CHEMISTRY (XI)

Chapter 6

Chemical Bonding

Short Questions

1. What is octet rule? Give two examples of compounds which deviatefrom it.

Ans: The tendency of an atom to attain a maximum of eight electrons in the valance shell is called octet rule.

There are many compounds in which atoms have no eight electrons (octet) in the valance shell after chemical combination. For example, BF_3, BCl_3, $AlCl_3$, SF_6, PF_5

2. Define chemical bond. Give examples.

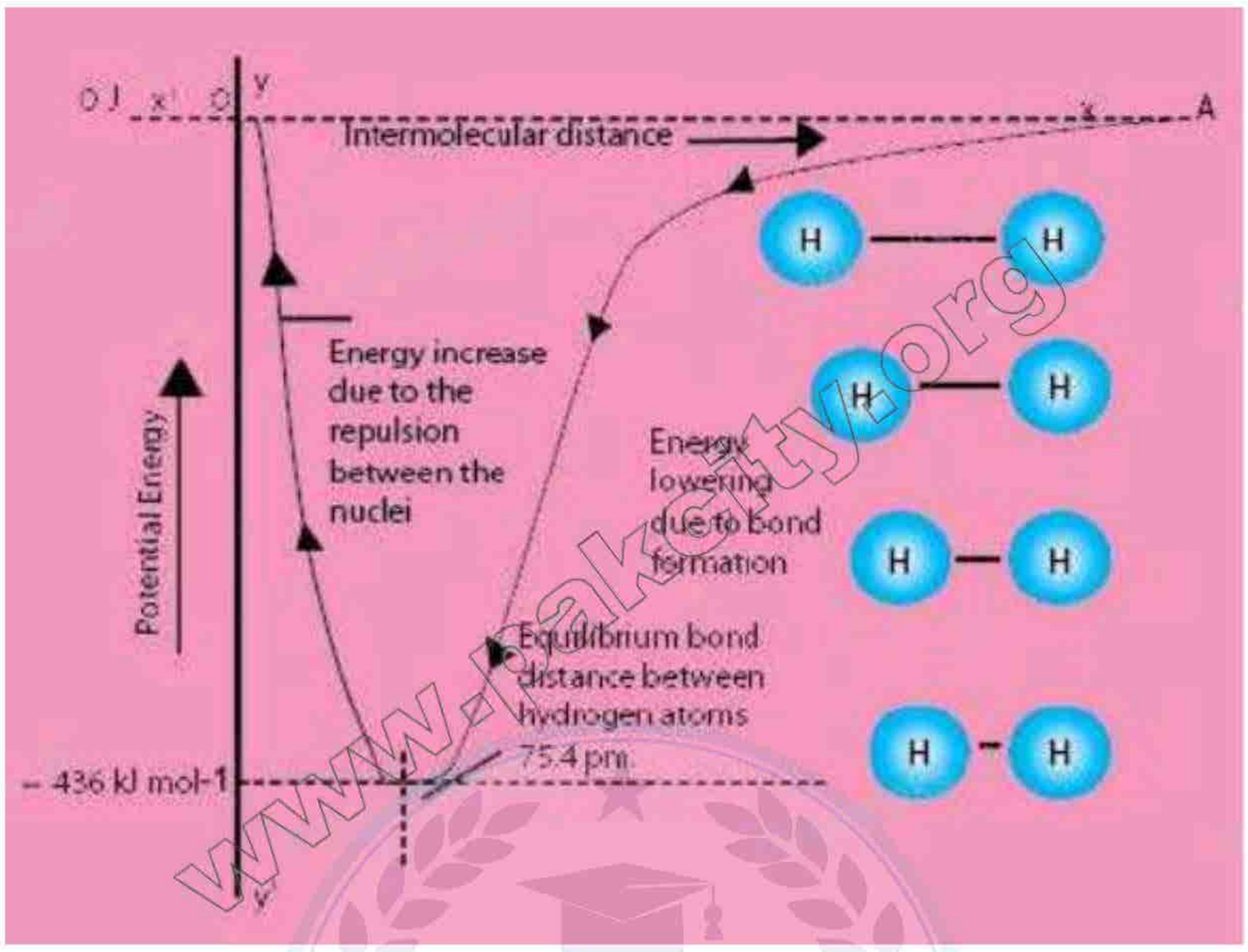
Ans: A chemical bond is the force which holds together two or more atoms or ions to forma large variety of compounds e.g. ionic bond, covalent bond.

3. Bond distance is the compromise distance. Justify.

Ans: When two atoms approach each other then there are two forces that operate i.e. attractive forces and repulsive forces. Attractive forces are between the nucleus of one atom and electrons of the other atom and repulsive forces are between the nuclei and electrons of both the atoms. There is competition between the attractive and repulsive forces till a point is reached where attractive forces dominate and repulsive forces are minimized. This is the point where bond is formed so bond distance is called the compromise distance between two atoms making up a molecule.

4. The compromise distance between the two bonded hydrogen atoms is 75.4 pm. Justify.

Ans: When two atoms come close to each other, they have attraction as well as repulsion between them. These bonded atoms stay at the distance, where the attraction is maximum. This distance is also called bond length. These bonded atoms cannot fly apart from each other in ordinary conditions. In H_2 molecule compromised distance between the two hydrogen atoms is 75.4pm, which is also called equilibrium bond distance or bond length of H_2 molecule.



5. Define atomic radius.

Ans: The atomic radius means the average distance between the nucleus of the atom and its outermost electronic shell.

6. Why atomic radii cannot be determined accurately?

Ans: The radius of an atom cannot be determined precisely due to the following reasons:

(i) There is no sharp boundary of an atom. The probability of finding an electron never becomes exactly zero even at large distances from the nucleus.

(ii) The electronic probability distribution is affected by neighbouring atoms. For this reason, the size of an atom may change from one compound to another.

7. Give trend of atomic radii in the periodic table.

Ans: Atomic radii decrease from left to right in the periodic table because the effective nuclear charge increases from left to right and shielding effect remains constant. It increases from top to bottom in the periodic table because the effective nuclear charge decreases and shielding effect increases.

8. Define ionic radius.

Ans: The ionic radius of an ion is the average distance between nucleus of ion and the outer shell. It is measured in pm or nm etc.

9. Why the radius of Cl^- ion increases from 99 pm to 181 pm OR Why the ionic radius is greater than atomic radius?

Ans: The radius of anion is larger than its original atom. The increase in the size of anion is due to the increase in electron-electron repulsion because of the increase in the valance shell electrons. We can say that effective nuclear charge decreases by the increase of one electron in the valance shell. This causes expansion of the shells. For example, size of chloride ion Cl^- is greater than Cl. Size of chlorine atom Cl is 99 pm while size of Cl^- is 181 pm.

10. No bond in chemistry is 100% ionic. Justify.

Ans: For a compound to be 100% ionic the electronegativity difference between the elements must be 4.0. So, no bond in chemistry is 100% ionic in nature. Highest ionic characters are present in CsF because Cesium is least electronegative and fluorine is most electronegative element. Difference of electronegativity is 4 – 0.7 = 3.3. Cesium fluoride is 92% ionic and has 8% covalent character in it. NaCl is 72% ionic and 28% covalent.

11. Define ionization energy. Give an example.

Ans: The ionization energy of an element is the minimum energy required to remove an electron from the outermost orbit of an isolated gaseous atom to form a positive ion. The process is called ionization, e.g.

$$Mg \rightarrow Mg^{+} + e^{-} \qquad \Delta H = 738 kJmol^{-1}$$

12. Give the trend of ionization energy in the periodic table.

Ans: Ionization energy increases from left to right in the periodic table because the atomic size decreases, shielding effect remains constant and effective nuclear charge increases. It decreases from top to bottom because the atomic size increases, shielding effect increases and effective nuclear charge decreases.

13. Mention the factors affecting ionization energy.

Ans: It is observed that the ionization energies of atoms depend upon the following factors:

(i) Atomic radius

(ii) Nuclear charge or proton number of the atom

(iii) Shielding effect of inner electrons

(iv) Nature of orbital

14. Ionization energy is an index to the metallic character. Justify.

Ans: Ionization energy is an index to the metallic character. The elements having low ionization energies are good metals and those having high ionization energies are non-metals. Those with intermediate values are mostly metalloids.

15. Define electron affinity. Give an example.

Ans: The electron affinity of an atom is the energy released or absorbed when one or more electrons are added to an empty or partially filled orbital of an isolated gaseous atom in its valence energy level to form negative ion. e.g.,

$$Cl(g) + e^- \rightarrow Cl^-(g) \qquad \Delta H = -349 kJmol^{-1}$$

16. Give the trend of electron affinity in the periodic table.

Ans: Electron affinities increase from left to right in period of the periodic table as the effective nuclear charge increases, shielding effect remains constant and size of atom decreases. There is more attraction of nucleus for outermost shell electrons, so electron affinity increases from left to right. Electron affinity decreases from top to bottom as the effective nuclear charge decreases, shielding effect increases and size of atom increases. There will be less attraction of nucleus for outermost shell electrons, so electron affinity decreases from top to bottom.

17. Define electronegativity. Give an example.

Ans: The tendency of an atom to attract a shared electron pair towards itself is called its electronegativity. For example, electronegativity of Hydrogen is 2.1 and that of Fluorine is 4.0.

18. Give the trend of electronegativity in the periodic table.

Ans: Electronegativity increases from left to right in the periodic table because atomic size decreases, shielding effect remains constant, effective nuclear charge increases. It decreases from top to bottom because atomic size increases, shielding effect increases and effective nuclear charge decreases.

19. Define ionic bond. Give an example.

Ans: According to the Lewis theory, ionic bond is formed by the complete transfer of one or more electrons from an atom with low ionization energy to another atom with high electron affinity. For example:

$$K(2,8,8,1) \rightarrow K^+(2,8,8) + e^- \qquad \Delta H = 419.0\,kJmol^{-1}$$

$$:\underset{\cdot\cdot}{\ddot{Cl}}\cdot + e^- \rightarrow :\underset{\cdot\cdot}{\ddot{Cl}}^- : \qquad \Delta H = -349\,kJmol^{-1}$$

20. Why is the radius of cation smaller than its parent atom?

Ans: Size of cation is smaller than its parent atom because of the imbalance of electron-proton ratio which can be further explained as:

i. No. of protons are greater than no. of electrons so nuclear charge increases

ii. Nucleus hold increases on the remaining electrons

iii. In some cases number of shells also decreases

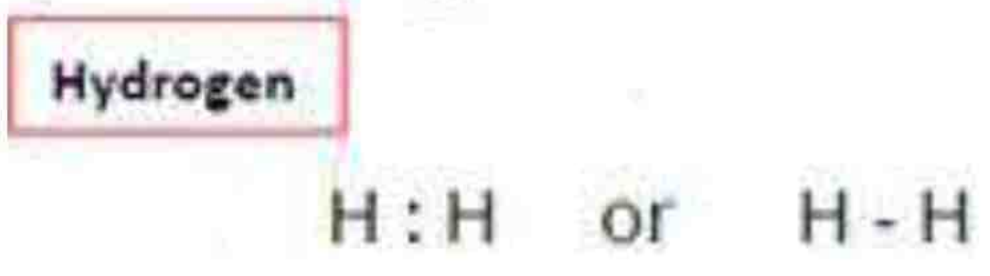
The radius of Na^+ ion is smaller than Na atom. The radius of Na^+ ion is 95 pm while Na is 186 pm. By losing one electron, the effective nuclear charge increases and shells shrink to smaller size.

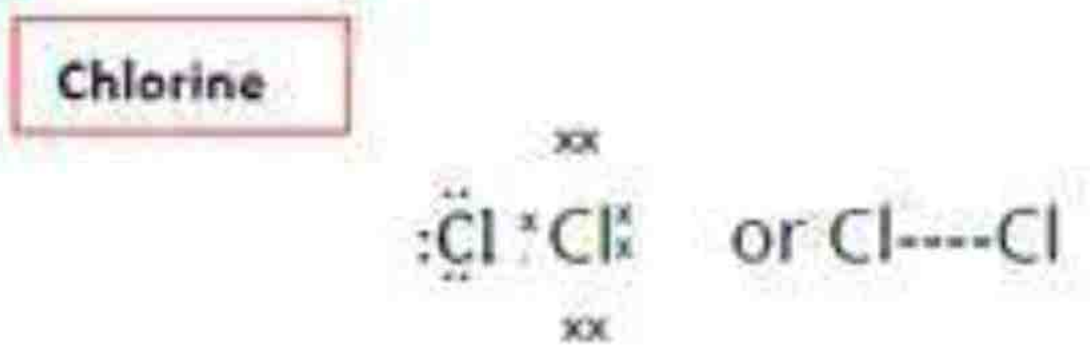
21. How electronegativity helps us to understand nature of bonds?

Ans: Criteria of electronegativity also helps us to understand the nature of bond. In order to decide the % of ionic nature in a compound, it is better to note the difference of electronegativity between the bonded atoms. If the difference is 1.7 or more then the bond is said to be ionic. Keeping this aspect in view, NaCl has 72% ionic character. CsF has 92% ionic character and calculations tell us that there is no bond with 100% ionic character.

22. Define covalent bond. Give an example.

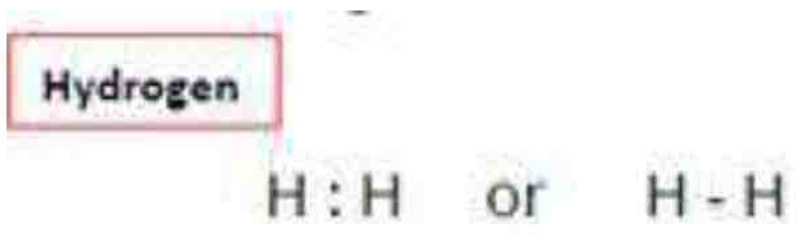
Ans: According to Lewis and Kossel, a covalent bond is formed by the mutual sharing of electrons between two atoms. For example:





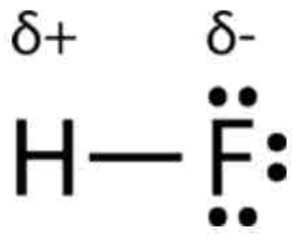
23. Define non-polar covalent bond. Give an example.

Ans: In such bonds, the bonding electron pairs are equally shared. For example, in H_2 or Cl_2 molecules, the two electrons forming the covalent bond are equally shared by the two identical atoms having same electronegativities. For example:

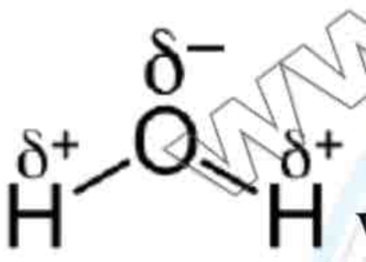


24. Define polar covalent bond. Give an example.

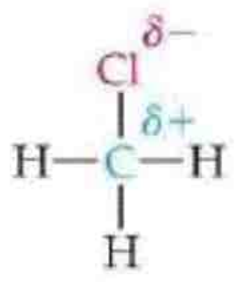
Ans: When two different atoms are joined by a covalent bond, the electron pair will be displaced towards the more electronegative atom and the pair shall go away from less electronegative atom. In this one atom gets partial negative charge and other partial positive charge. Such bond is considered polar. For example,



Hydrogen fluoride



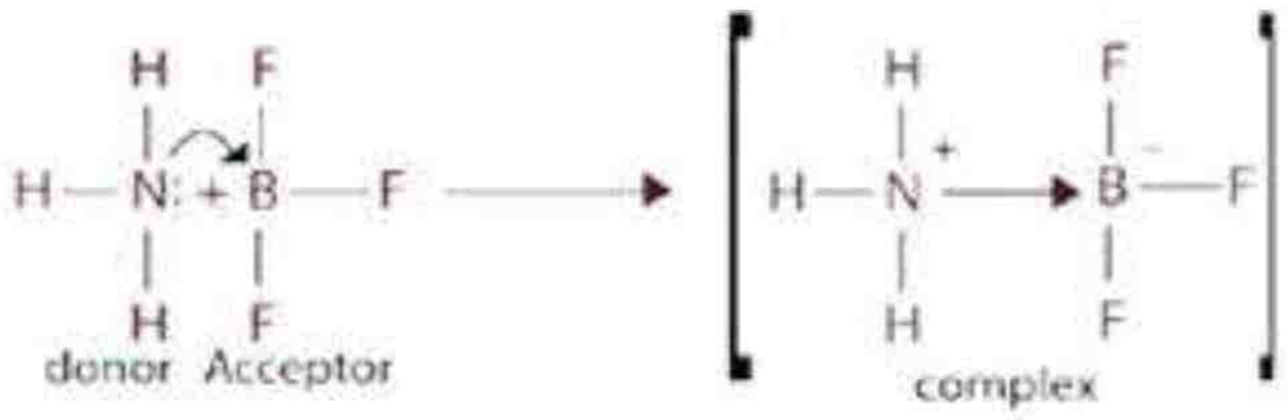
Water



Methyl chloride

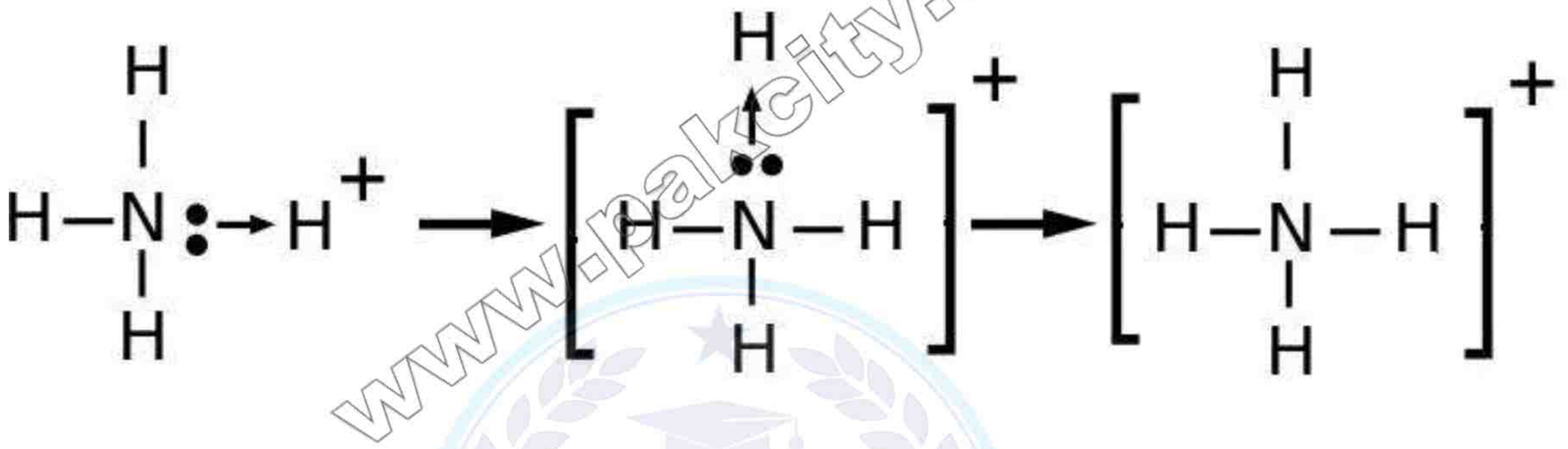
25. Define coordinate covalent bond. Give an example.

Ans: A coordinate covalent bond is formed between two atoms when the shared pair of electrons is donated by one of the bonded atoms. For example, bond formation between NH_3 and BF_3.



26. Distinction between covalent and coordinate covalent bond vanishes after bond formation. Justify.

Ans: Coordinate covalent bond is just like a single bond. For example, this bond is formed by the empty orbital of H^+ and completely filled orbital of Nitrogen. According to Lewis concept, nitrogen donates a pair of electron to H^+, So NH_3 is a donor while H^+ is acceptor. Usually, this bond is donated by an arrow. Arrow head is from donor towards acceptor.



Experimentally all four N-H bonds are identical in bond length and bond strength. Therefore, ammonium ion is written as NH_4^+ and it is that after bond formation the distinction between coordinate covalent bond and covalent bond vanishes.

27. Differentiate between covalent bond and coordinate covalent bond.

Ans:

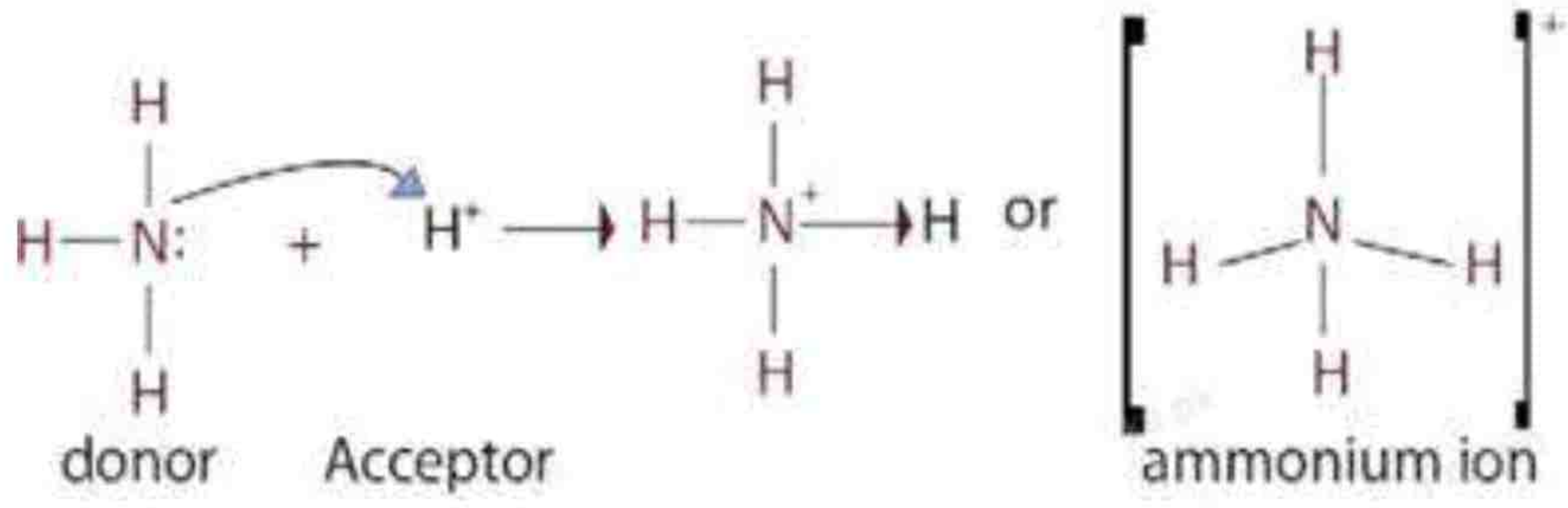
Covalent Bond	Coordinate Covalent Bond
The bond formed by the mutual sharing of electrons between two atoms is called covalent bonds.	The bond formed when the shared pair of electrons is donated by one of the bonded atoms is called coordinate covalent bond.
Example: CH_4, H_2O	Example: NH_4^+, BF_4^-
Two atoms donate an equal share of electrons.	One atom called donor donates a pair of electrons and one atom called acceptor accepts the pair of electron.

28. What are oxonium ions? Give an example.

Ans: All the alcohols and ethers offer their lone pairs to H^+ just like water to give coordinate covalent bonds. The ions so produced with positively charged oxygen atom are called oxonium ions.

29. How NH_4^+ ion is formed?

Ans: Ammonia donates its electron pair to H^+ ion to give NH_4^+ ion. All the four bonds behave alike in NH_4^+ ion.



30. What is the basic assumption of VSEPR theory?

Ans: The valence electron pairs (lone pairs and the bond pairs) are arranged around the central atom to remain at a maximum distance apart to keep repulsions at a minimum.

31. Why the non-bonding electron pairs occupy more space than the bonding electrons?

Ans: A bonding electron pair is attracted by both nuclei of atoms while non- bonding by only one nucleus. Because a lone pair experiences less nuclear attraction, its electronic charge is spread out more in space than that for bonding pair. As a result, the non- bonding electron pairs exert greater repulsive forces on bonding electron pairs and thus tend to compress the bond pairs.

32. Mention the magnitude of repulsions between the electron pairs in a given molecule.

Ans: The magnitude of repulsions between the electron pairs of a molecule decreases in the following order:

Lone pair- lone pair > lone pair -bond pair > bond pair - bond pair

These repulsions are called Van der Waal's repulsions.

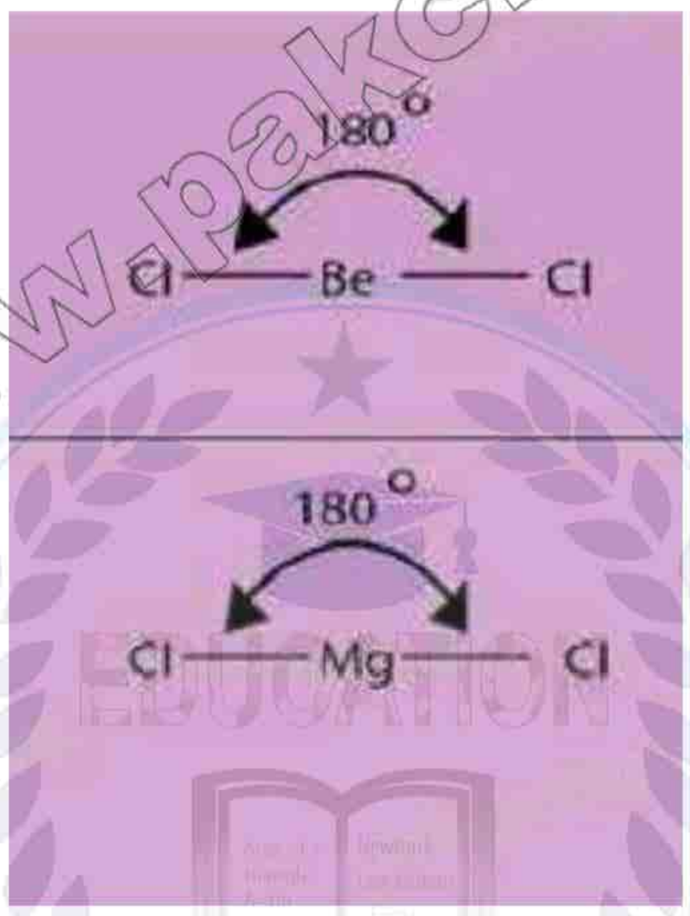
33. Why multiple bonds are considered as single bond in VSEPR theory?

Ans: The two electron pairs of a double bond and three electron pairs of a triple bond, contain a higher electronic charge density. Therefore, they occupy more space than one electron pair of a single bond but behave like a single electron pair in determining the geometry of the molecule. This is because they tend to occupy the same region between the two nuclei like a single bond.

34. Explain AB_2 type molecules according to VSEPR theory.

Ans: In such molecules two electron pairs around the central atom are arranged at an angle of 180° in order to minimize repulsions between them. Thus, they form a linear geometry.

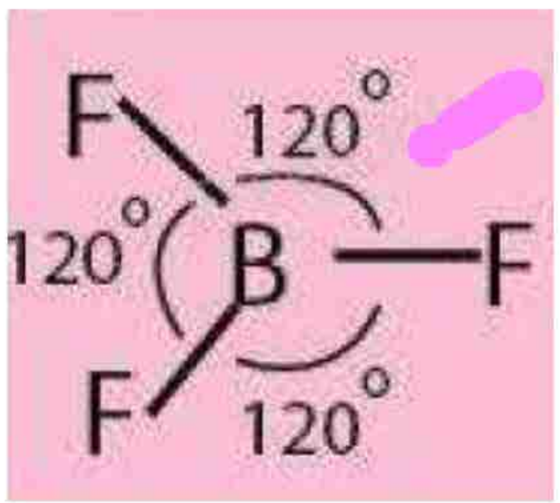
Beryllium chloride is a typical linear molecule which contains two electron pairs. $MgCl_2$, $CaCl_2$, $SrCl_2$, $CdCl_2$ and $HgCl_2$ are also linear molecules.



35. Explain AB_3 type molecules with no lone pairs.

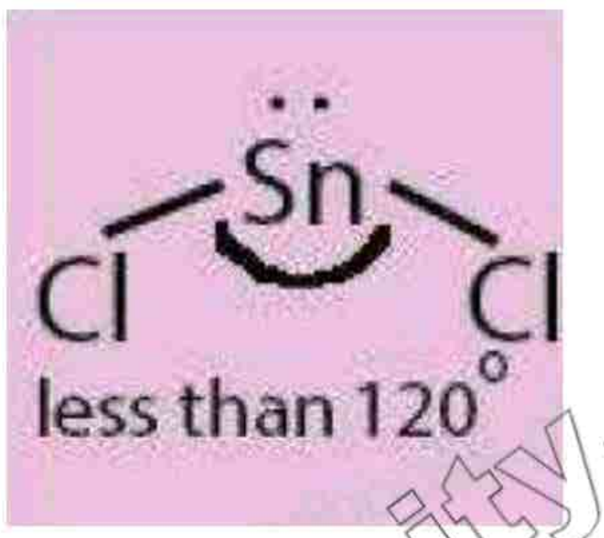
Ans: In such molecules, central atom contains three bonding electron pairs, which are arranged at maximum distance apart at a mutual angle of 120°, giving a triangular planar geometry. The boron atom in BH_3 is surrounded by three charge clouds, which remain farthest apart in one plane, each pointing towards the corners of an equilateral triangle. Thus, BH_3, molecules has a trigonal planar geometry, with each H- B-H bond angles of 120°.

Similar geometries are of hydrides of group III-A (AlH_3, GaH_3, InH_3 and TlH_3)and their halides (BF_3, $AlCl_3$,etc.)



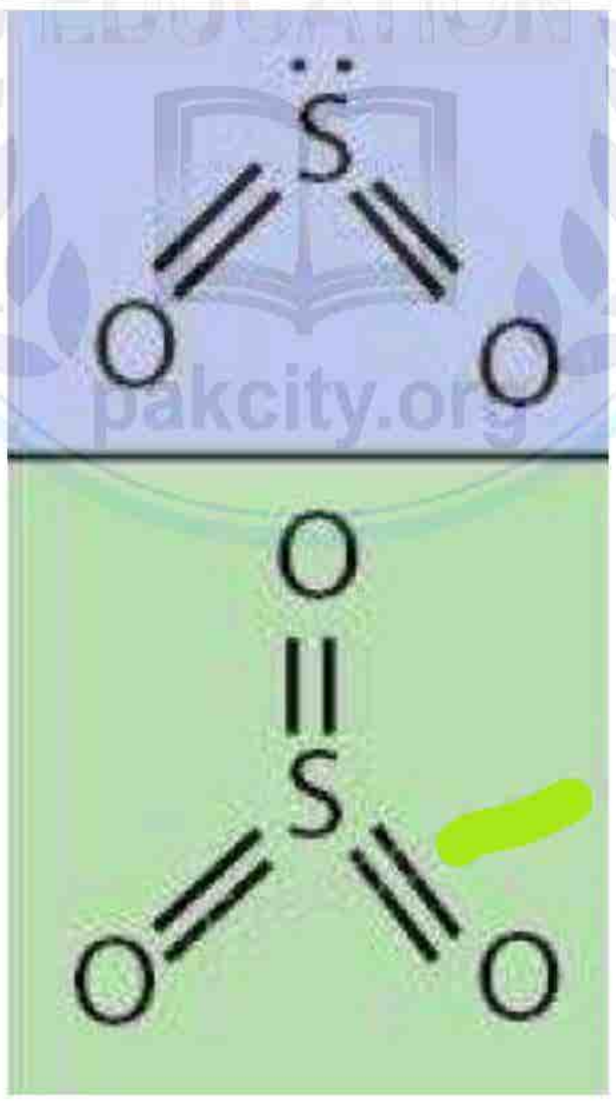
36. Explain AB_3 type molecules with one lone pair and two bond pairs.

Ans: In $SnCl_2$, one of the corners of the triangle is occupied by a lone pair, giving rise to a distorted triangular structure in vapor phase.



37. Explain AB_3 type molecules with multiple bonds.

Ans: In SO_2, one corner of triangle is occupied by a lone pair and two corners each by S=O double bond, while in SO_3 all three regions, each are occupied by S = O bonds. This structure of SO_3 is perfectly triangular.



38. Explain AB_4 type molecules with no lone pairs.

Ans: The charge clouds due to four electron pairs avoid their electrostatic repulsions by drifting apart to maintain a mutual bond angle of 109.5°. These molecules have shape of a regular tetrahedron.

Example

Each of the four valence electrons of carbon pair up with one electron of hydrogen in methane.

$${}_6C = 1s^2, 2s^1, 2px^1, 2py^1, 2pz^1$$

The four electron pairs are directed from the center towards the corners of a regular tetrahedron, with each apex representing a hydrogen nucleus. The arrangement permits a non-planar arrangement of electron pairs. Each H-C-H bond is perfectly 109.5 °. Molecules of similar geometry are SiH_4, GeH_4, CCl_4.

39. Explain AB_4 type molecules with one lone pair and three bond pairs.

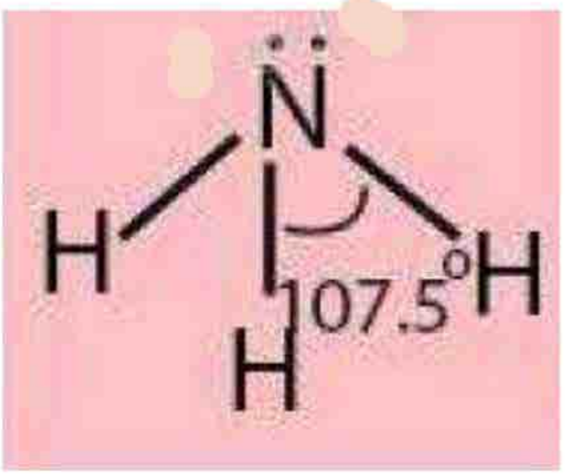
Ans: The charge cloud of lone pair electrons (non-bonding electrons) spreadsout more than that of bonding electrons.

Large lone pair charge cloud tends to compress the bond angles in restof the molecules.

Ammonia, NH_3 is a typical example.

$${}_7N = 1s^2, 2s^2, 2px^1, 2py^1, 2pz^1$$

The non-bonding electron in 2s orbital takes up more space and exerts a strong repulsive force on the bonding electron pairs. Consequently, to avoid a larger repulsion, the bonding electron pairs move closer that reduces the ideal bond angle from 109.5°to 107.5°. The resultant molecule has triangular pyramidal geometry instead of tetrahedral.



40. Why the bond angle is reduced to 102° in NF_3?

Ans: Substitution of hydrogen in NH_3 with electronegative atoms like F or Cl further reduces the bond angle. In NF_3, the strong polarity of N-F bond pulls the lone pair of N atom closer to its nucleus. It exerts a stronger repulsion over bonding electrons. The angle further shrinks to 102°. Moreover, the bond pairs N-F bonds are closer to F atoms than N atoms. The increased distances in these bond pairs make their repulsions less operative.

41. Explain AB_4 type molecules with two lone pairs and two bond pairs.

Ans: Presence of two lone pairs, introduces three types of repulsion i.e., lone pair-lone pair, lone pair-bond pair and bond pair-bond pair repulsion.

For example: water (H_2O), a triatomic molecule is expected to be an AB_2 type linear molecule like $BeCl_2$ and CO_2. But experimental evidence confirms a bent or angular geometry. VSEPR theory explain geometry of water molecule.

$${}_8O = 1s^2, 2s^2, 2px^2, 2py^1, 2pz^1$$

Two of the corners of a tetrahedron are occupied by each of the two lone pairs and remaining by bond pairs. By lone pair-lone pair and lone pair-bond pair repulsions the bond angle is further reduced to 104.5°.

42. Write two points of Valence bond theory.

Ans: According to valence bond theory,

i. The partially filled atomic orbitals overlap to form bonds but the individual character of atomic orbitals is retained.

ii. Greater the overlap, stronger will be the bond formed.

43. How hydrogen molecule is formed according to VBT?

Ans: As the two atoms of hydrogen approach each other, their 1s orbitals overlap, thereby giving the H-H bond. The electron density becomes concentrated between the two nuclei. The bond is called a sigma (σ) bond.

44. Define sigma bond.

Ans: A single bond is formed when two partially filled atomic orbitals overlap in such a way that the probability of finding the electron is maximum around the line joining the two nuclei.

45. How hydrogen fluoride molecule is formed according to VBT?

Ans: The H-F bond is formed by the pairing of electrons - one from hydrogen and one from fluorine.

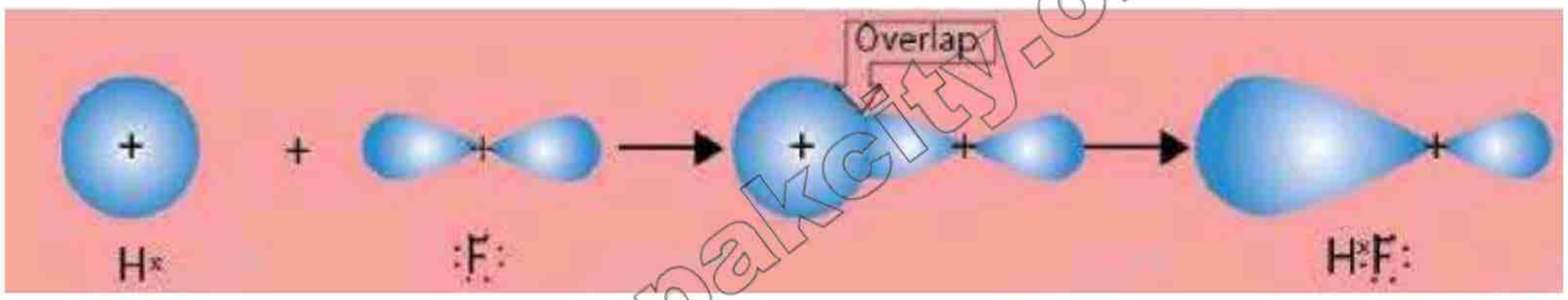
$$_{1}H = \overset{\boxed{\uparrow}}{1s}$$

$$_{9}F = \overset{\boxed{\uparrow\downarrow}}{2s}\,\overset{\boxed{\uparrow\downarrow}}{2p_x}\,\overset{\boxed{\uparrow\downarrow}}{2p_y}\,\overset{\boxed{\uparrow}}{2p_z}$$

The fluorine atom completes its 2p sub-shell by acquiring a share of an electron from hydrogen as shown below.

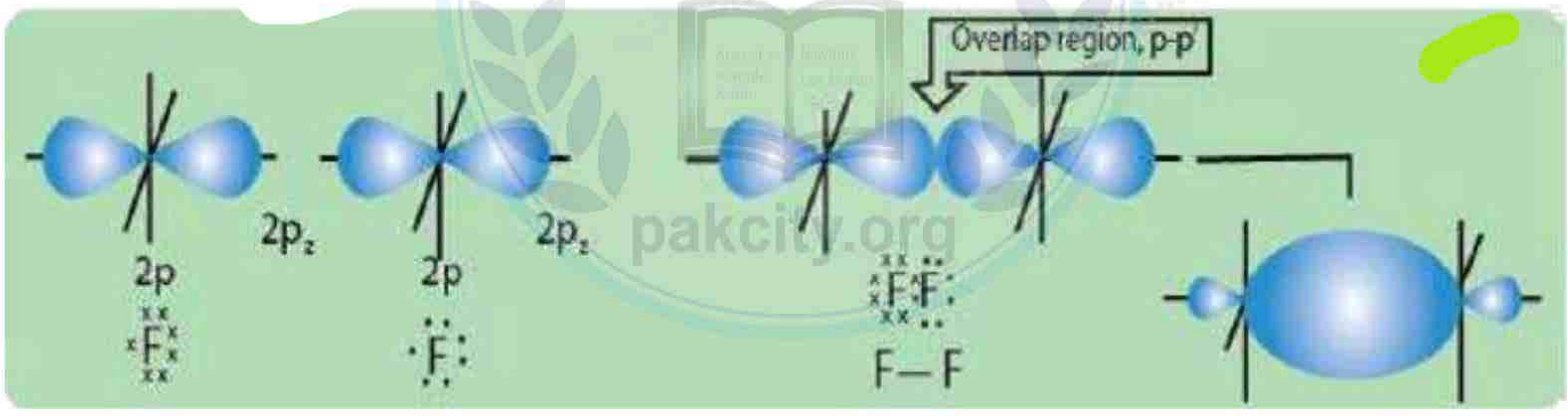
$$_{9}F = \overset{\boxed{\uparrow\downarrow}}{2s}\,\overset{\boxed{\uparrow\downarrow}}{2p_x}\,\overset{\boxed{\uparrow\downarrow}}{2p_y}\,\overset{\boxed{\uparrow\downarrow}}{2p_z}$$

There are then two orbitals plus two electrons whose spins can adjust so they are paired.



46. How fluorine molecule is formed according to VBT?

Ans: The sigma bond in the fluorine molecule, F_2 is formed by the overlap of half-filled 2pz orbital on each fluorine atom.

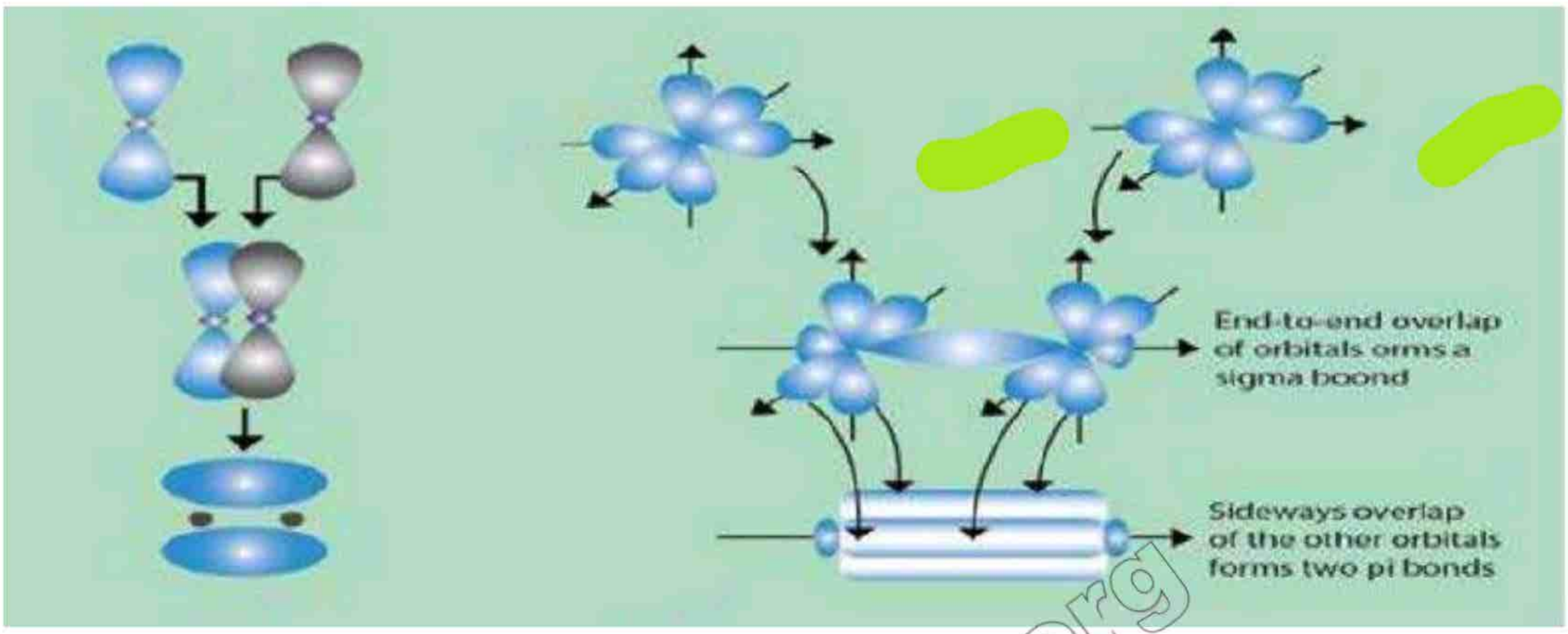


47. How nitrogen molecule is formed according to VBT?

Ans: Nitrogen atoms have the following electronic configuration

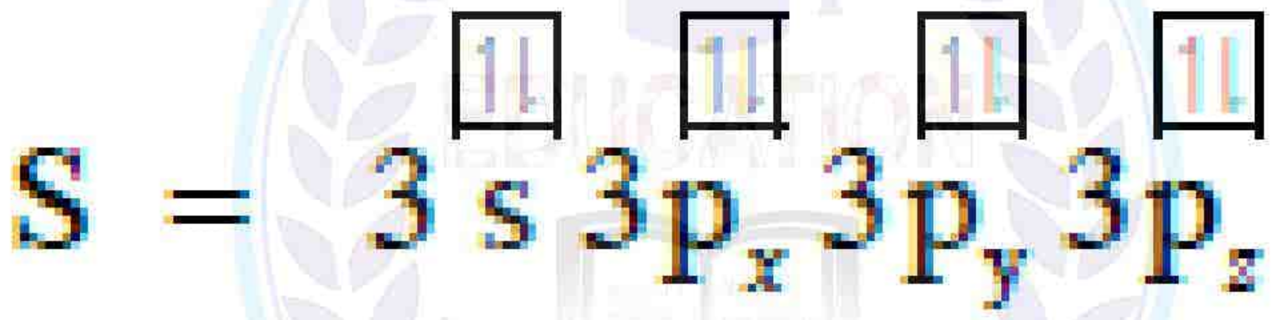
$$1s^2\ 2s^2\ 2px^1 2py^1 2pz^1$$

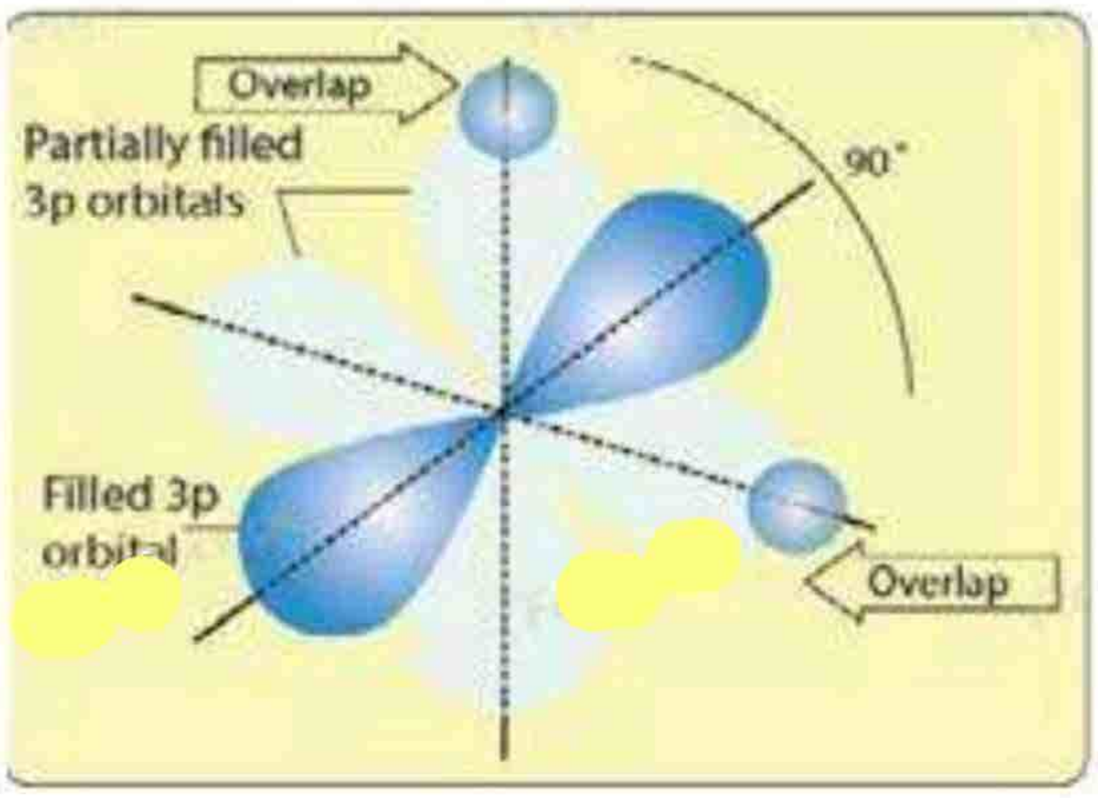
One end-to-end p orbital overlap occurs resulting in a sigma bond, the other two p orbital are aligned parallel to thecorresponding orbital in the other atom giving two pi bonds.

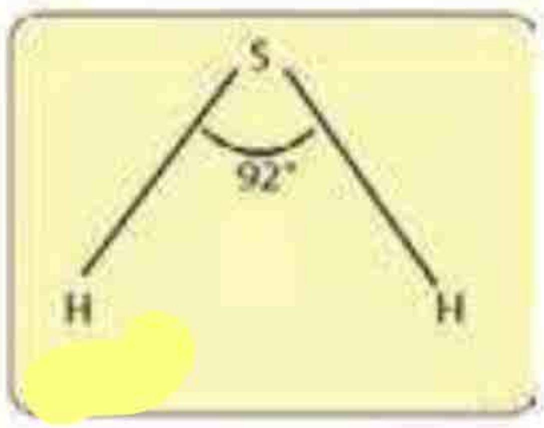


48. How H_2S molecule is formed according to VBT?

Ans: H_2S is a non-linear molecule, and the bond angle between the two H-S bonds is about 92°. Each two 3p orbitals of sulphur containing one electron can overlap with the 1s orbitals of hydrogen atoms.







49. Define hybridization.

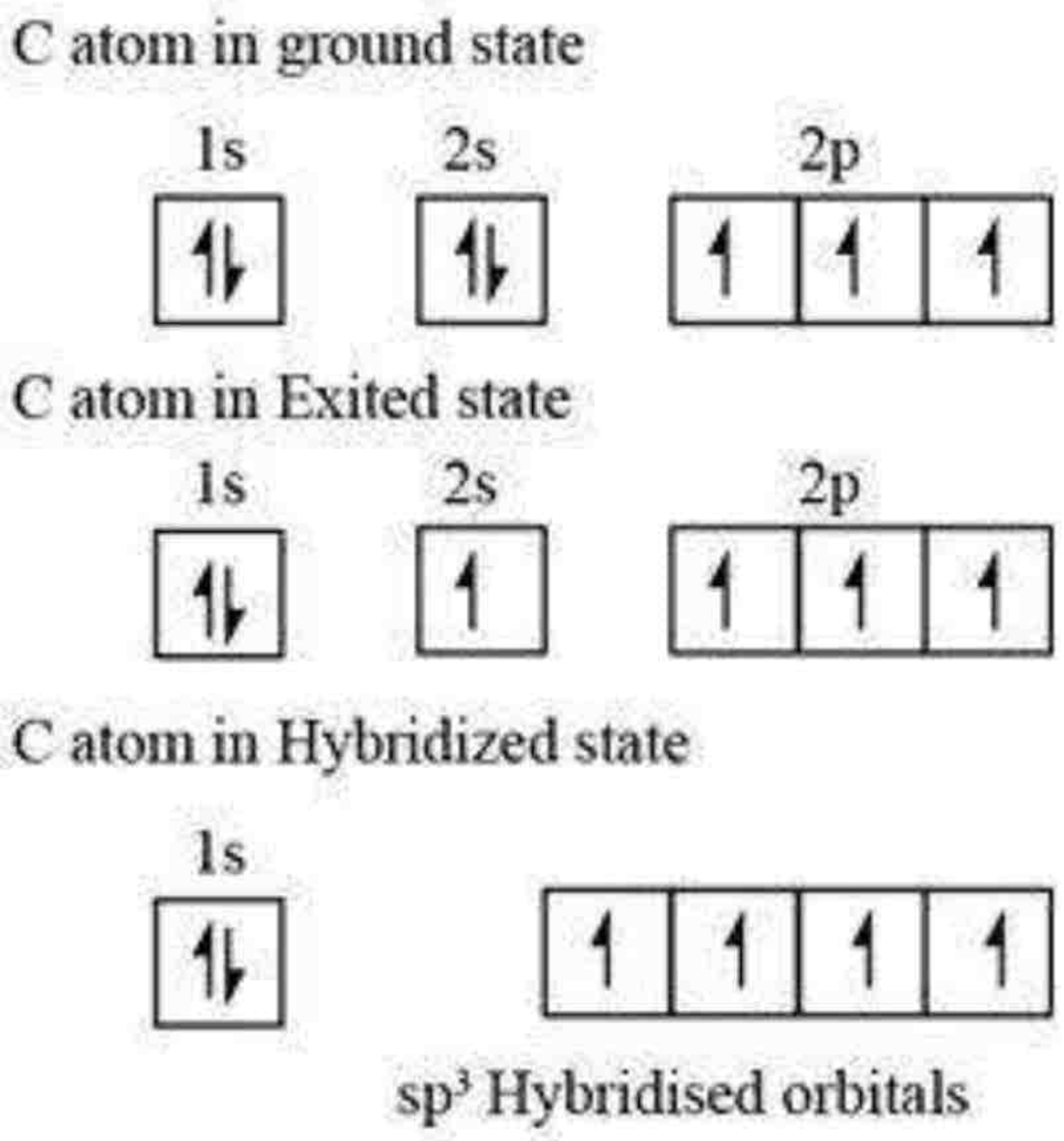
Ans: According to this, atomic orbitals differing slightly in energy intermix to form new orbitals of exactly same energies which are called hybrid atomic orbitals. They differ from the parent atomic orbitals in shape and possess specific geometry.

50. Define sp^3 hybridization. Give examples.

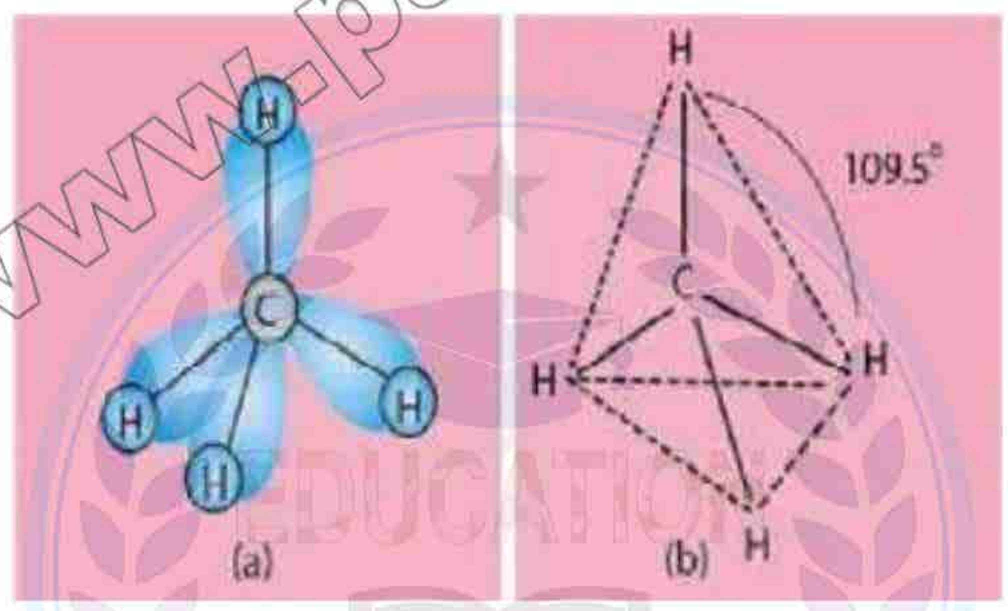
Ans: In sp^3 hybridization, one s and three p atomic orbitals of an atom intermix to form four equivalent orbitals called sp^3 hybrid atomic orbitals. Examples CH_4, NH_3 and H_2O.

51. Discuss structure of methane according to sp^3 (Similar question may beasked for ammonia and water)

Ans: Electronic configuration of $_6C$, its electronic excitation and hybridization is given as follows:



The four equivalent hybrid orbitals are directed towards the four corners of a regular tetrahedron. The hybrid orbitals are oriented in space in such a manner that the angle between them is 109.5°. They are six in number. Methane molecule is formed by the overlap of sp^3 hybrid orbitals of carbon with 1s orbitals of four hydrogen atoms separately to form four sigma bonds. The tetrahedral structure of CH_4 has four faces, four corners and six edges.



52. The bond angles of H_2O and NH_3 are not 109.5° like that of CH_4 although Oxygen and Nitrogen atoms are sp^3 hybridized. Why?

Ans: According to VSEPR theory, lone pairs occupy more space than bond pairs and cause more repulsions.

- H_2O has two lone pairs, so it repels the bond pairs much more and makes bond angle shorter till 104.5°.

- NH_3 has one lone pair that repels the three bond pairs so the bond angle between hydrogen atoms of ammonia is 107.5°.

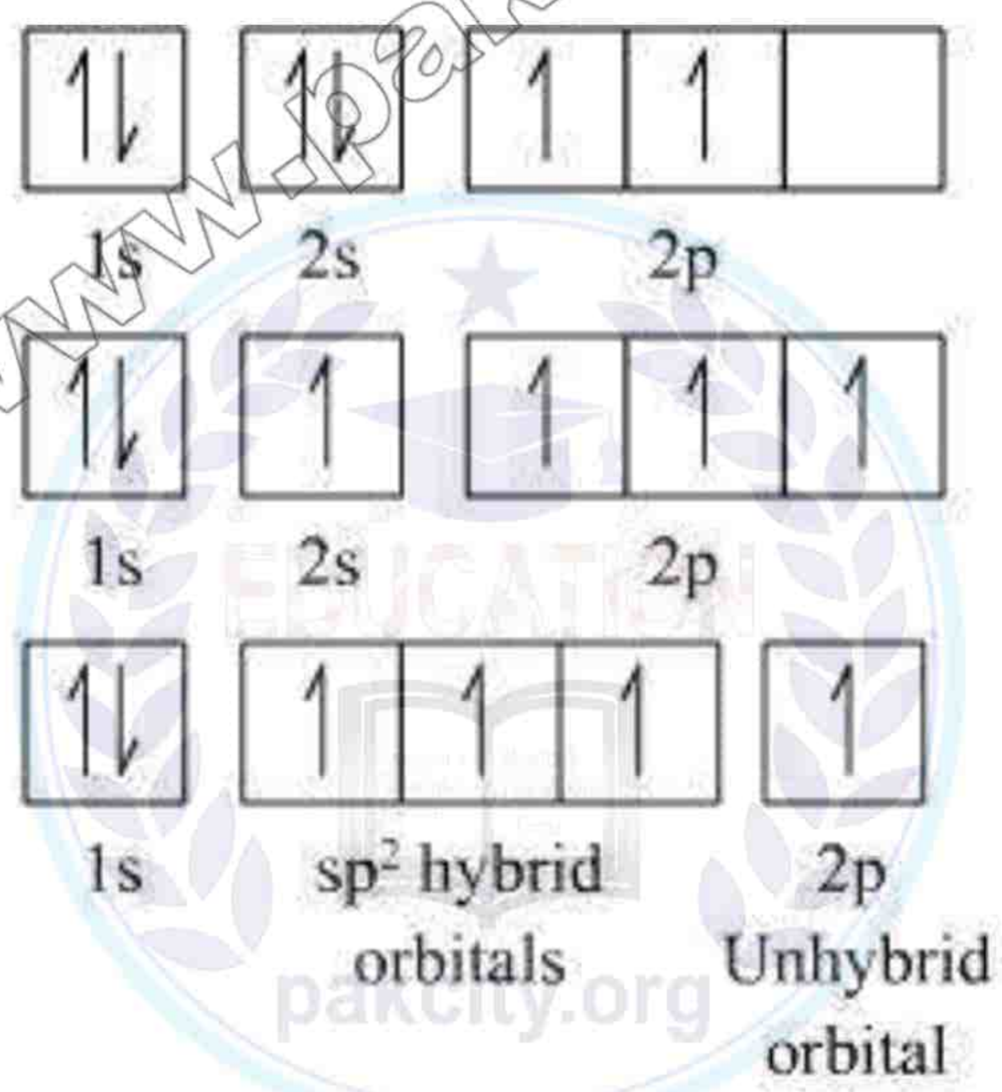
- CH_4 molecule has no lone pair, and each bond pair repels each other with equal force and bond angle between two adjacent hydrogen atoms becomes 109.5°.

53. Define sp^2 hybridization. Give examples.

Ans: In sp^2 hybridization, one 's' and two 'p' atomic orbitals of an atom intermix to give three orbitals called sp^2 hybrid orbitals. Example, BF_3 and Ethene.

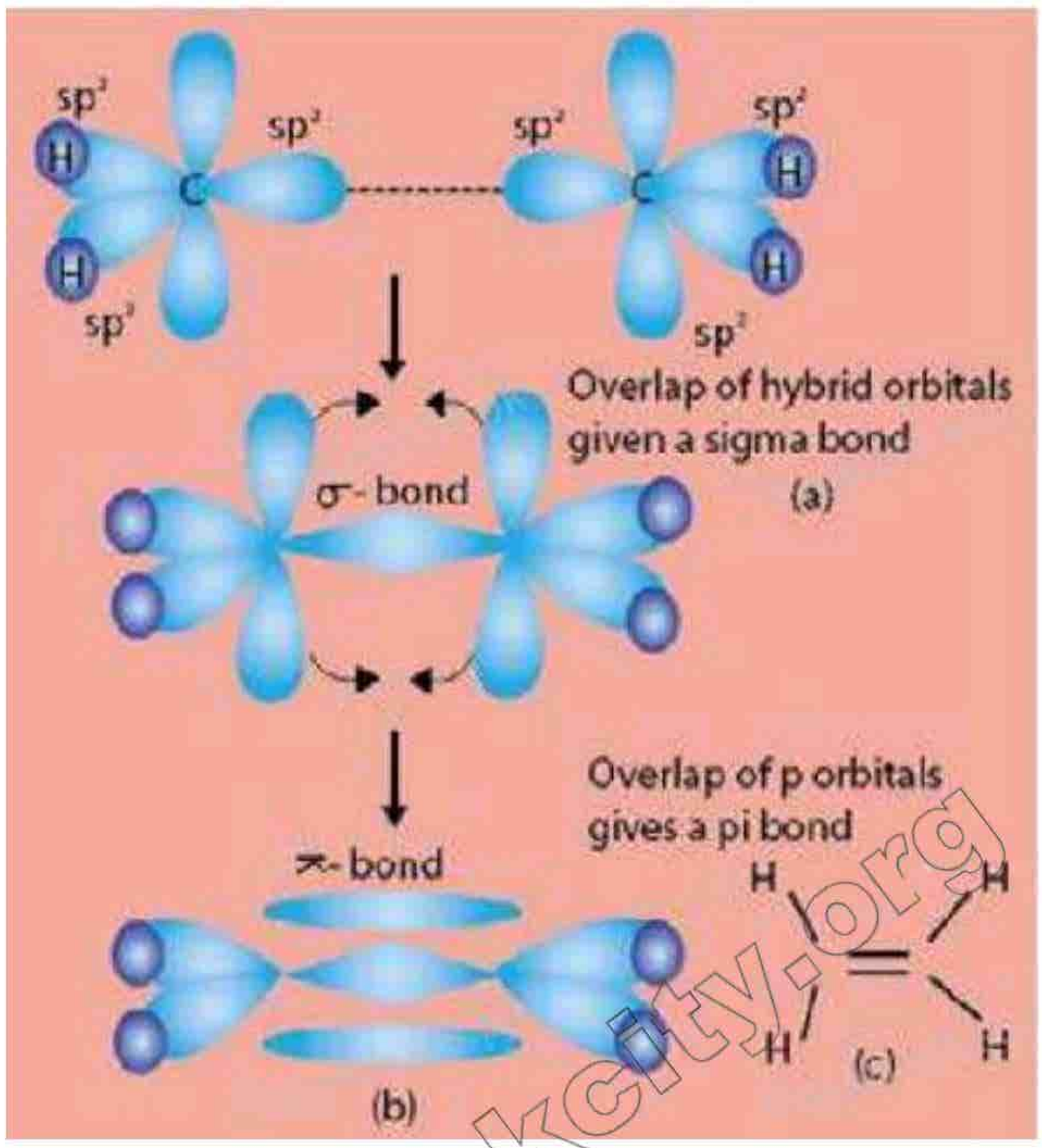
54. Explain structure of ethene (ethylene) according to hybridization.

Ans: Electronic configuration of $_6C$ is



Each carbon atom undergoes sp^2 hybridization to form three hybrid orbitals which are co-planar and are oriented at an angle of 120°.

One of the p-orbitals does not take part in hybridization. Each carbon atom undergoes sp^2-s overlaps with two hydrogen atoms and sp^2-sp^2 overlap between themselves to form sigma bonds. The partially filled p-orbitals undergo sideways overlap to form a π−bond.

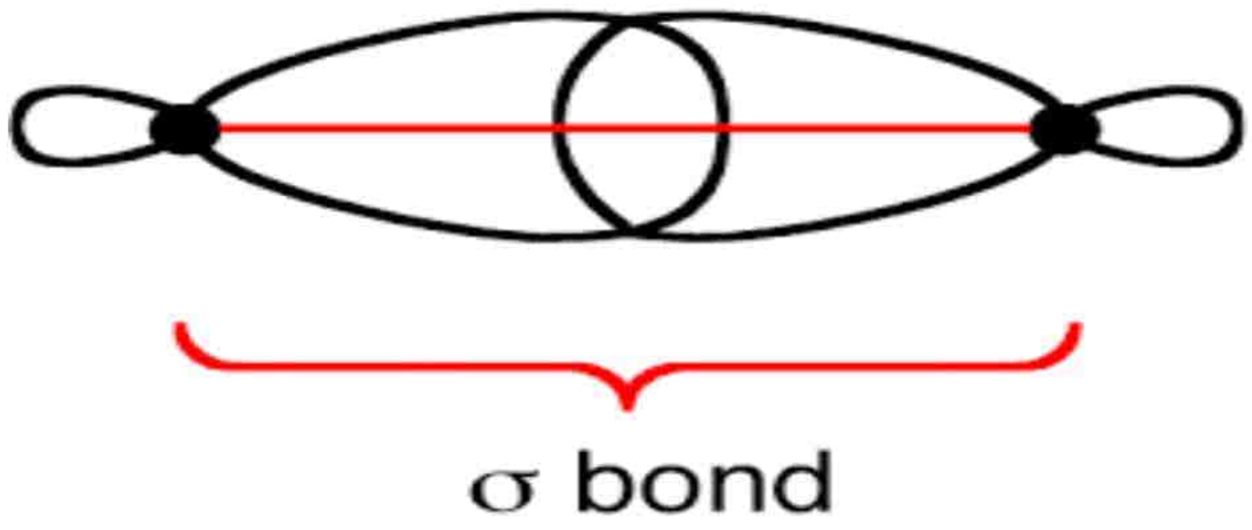


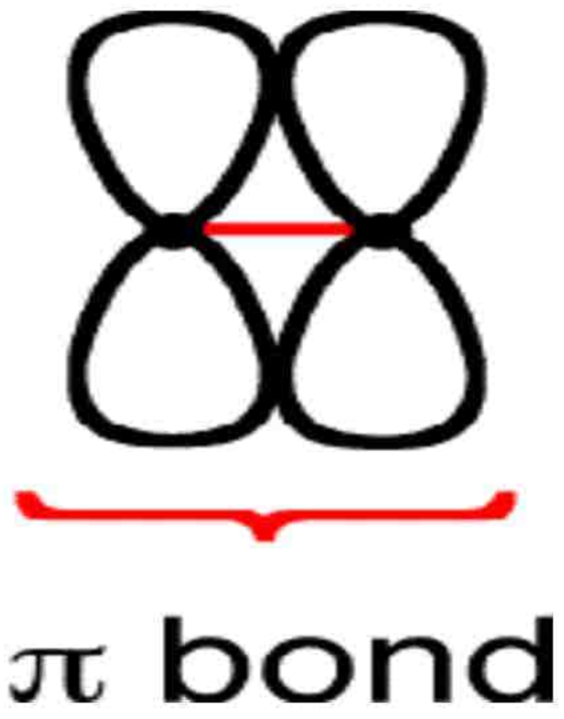
55. Define pi bond.

Ans: A pi-bond is formed by the sideways overlap of two half-filled co-planar p-orbitals in such a way that the probability of finding the electron is above or below the line joining the two nuclei.

56. π bonds are more diffused than σ bonds. Why?

Ans: In sigma bonded atoms, the electron density is between two nuclei or at bond axis. In pi bond, electronic cloud lies above and below the line joining the nuclei, so pi bonds are more diffused (spread) than sigma bonds.



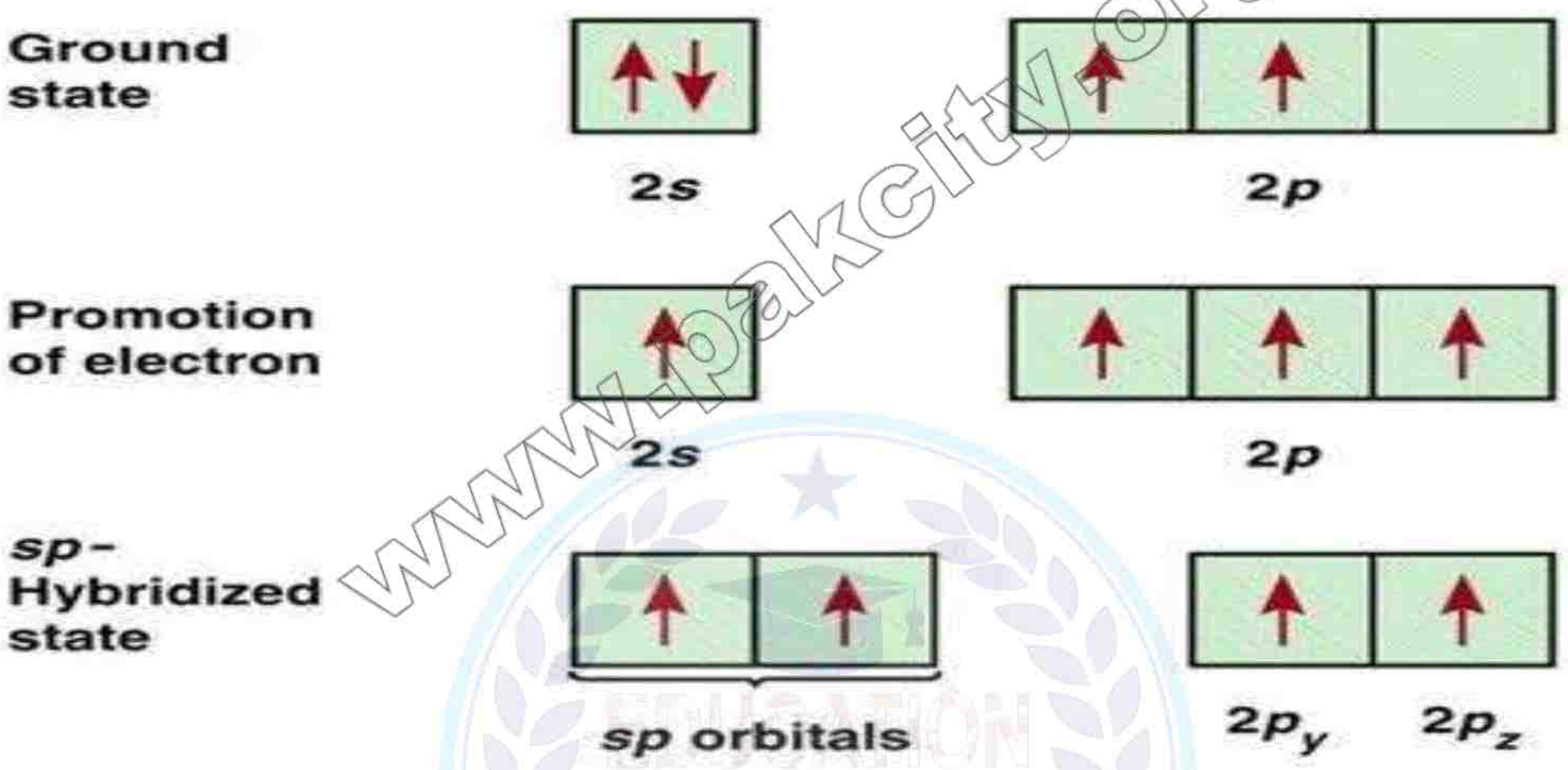


57. Define sp hybridization.

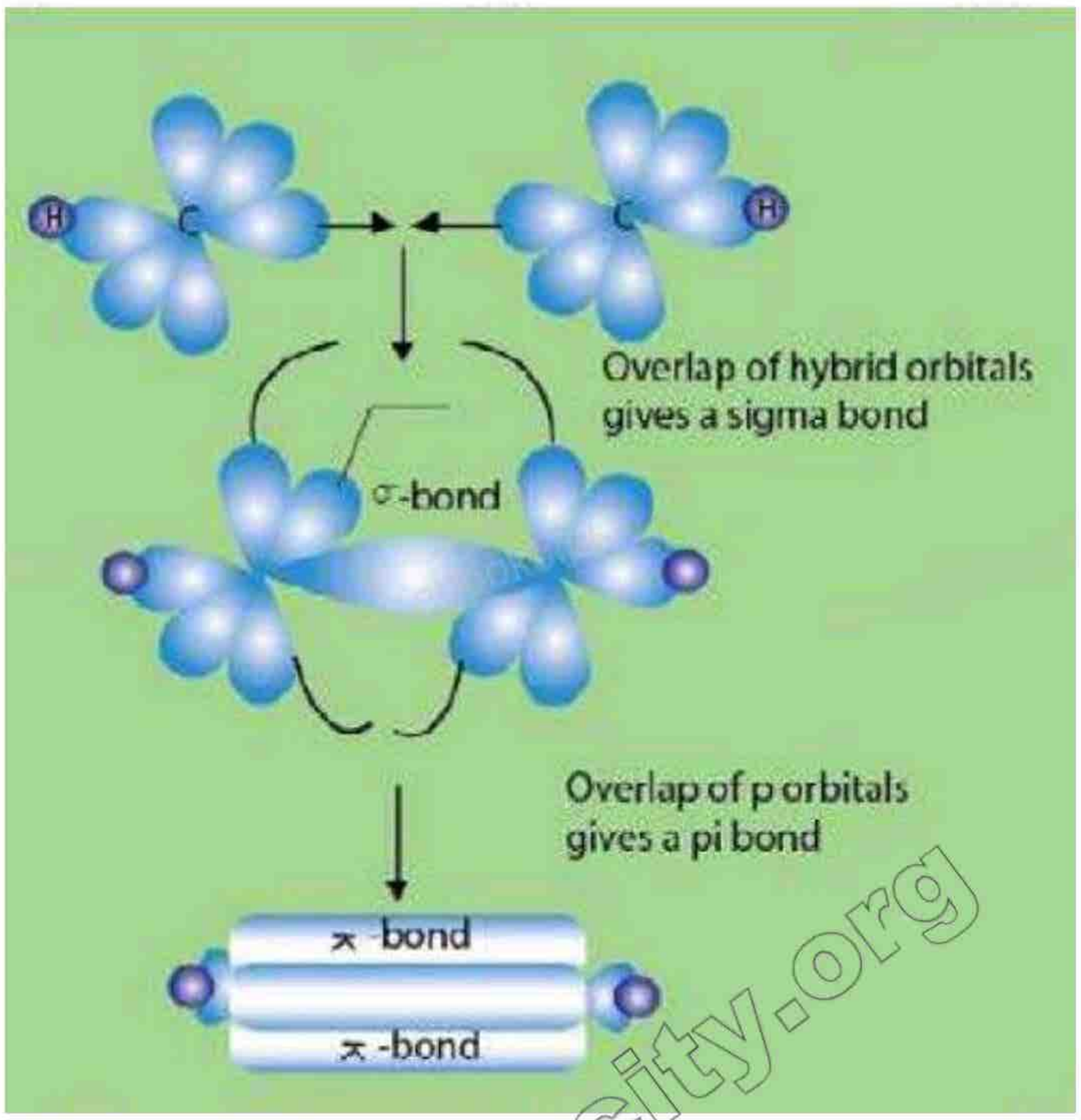
Ans: In sp hybridization, one ‘s’ and one ‘p’ orbitals intermix to form two sp-hybrid orbital. Examples, $BeCl_2$ and ethyne.

58. Explain structure of ethyne according to hybridization.

Ans: Electronic configuration of $_6C$ is



Each carbon atom undergoes sp-s overlap with one hydrogen atom and sp-sp overlap with other carbon atom. Each carbon atom is left with two unhybridized p orbitals perpendicular to the plane of sp hybrid orbitals. The sideways p overlap between the p-orbitals results in the formation of two pi bonds. Ethyne molecule contains one σ and two pi bonds between the two carbon atoms and each carbon atom is bonded with one H atom through σ bond.



59. What is the difference between VBT and MOT?

Ans: Following are the points of difference between VBT and MOT:

1. VBT considers overlap of valence shells whereas MOT considers atoms as a whole to participate in bonding.

2. Simplest diagrammatic representation of no bond formation was given by MOT not VBT.

3. MOT very well explains the paramagnetism and diamagnetism phenomena in molecules whereas VBT fails to explain it.

4. Simplest formula to calculate bond order was given by MOT not VBT.

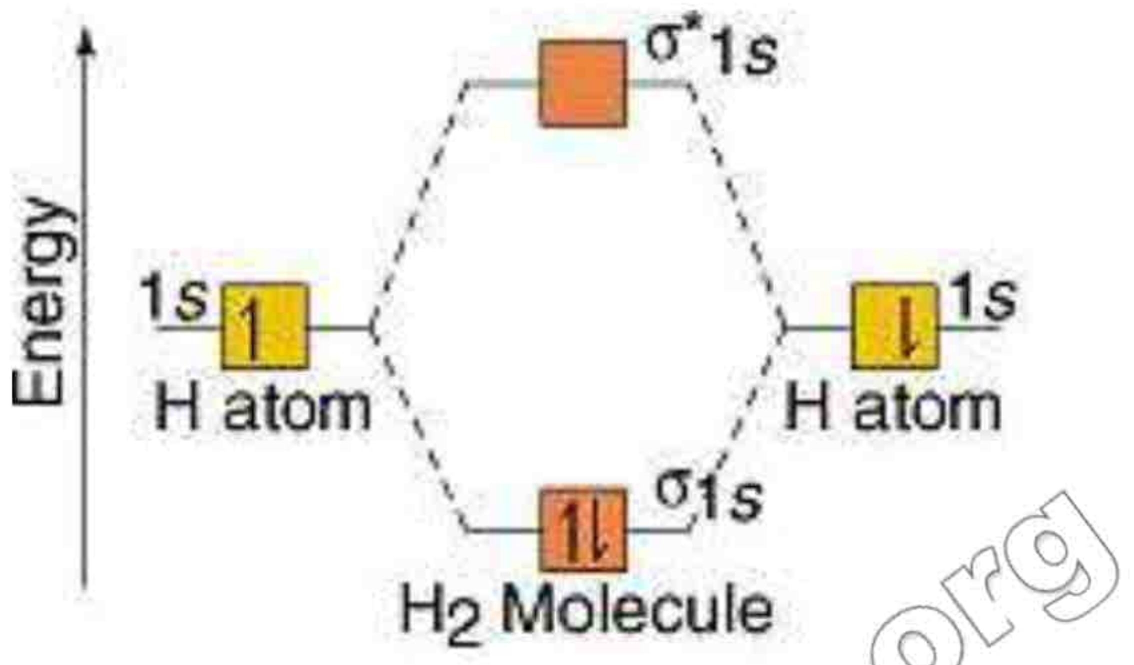
60. Define bond order. Give an example.

Ans: The number of bonds formed between two atoms after the atomic orbitals overlap is called the bond order and is taken as half of the difference between the number of bonding electrons and anti-bonding electrons. The number of bonds formed between H-atoms in hydrogen molecule may be calculated as follows:

Number of electrons in the bonding orbitals = 2

Number of electrons in the anti-bonding orbitals =2

$$\text{Bond order} = \frac{2-0}{2} = 1$$

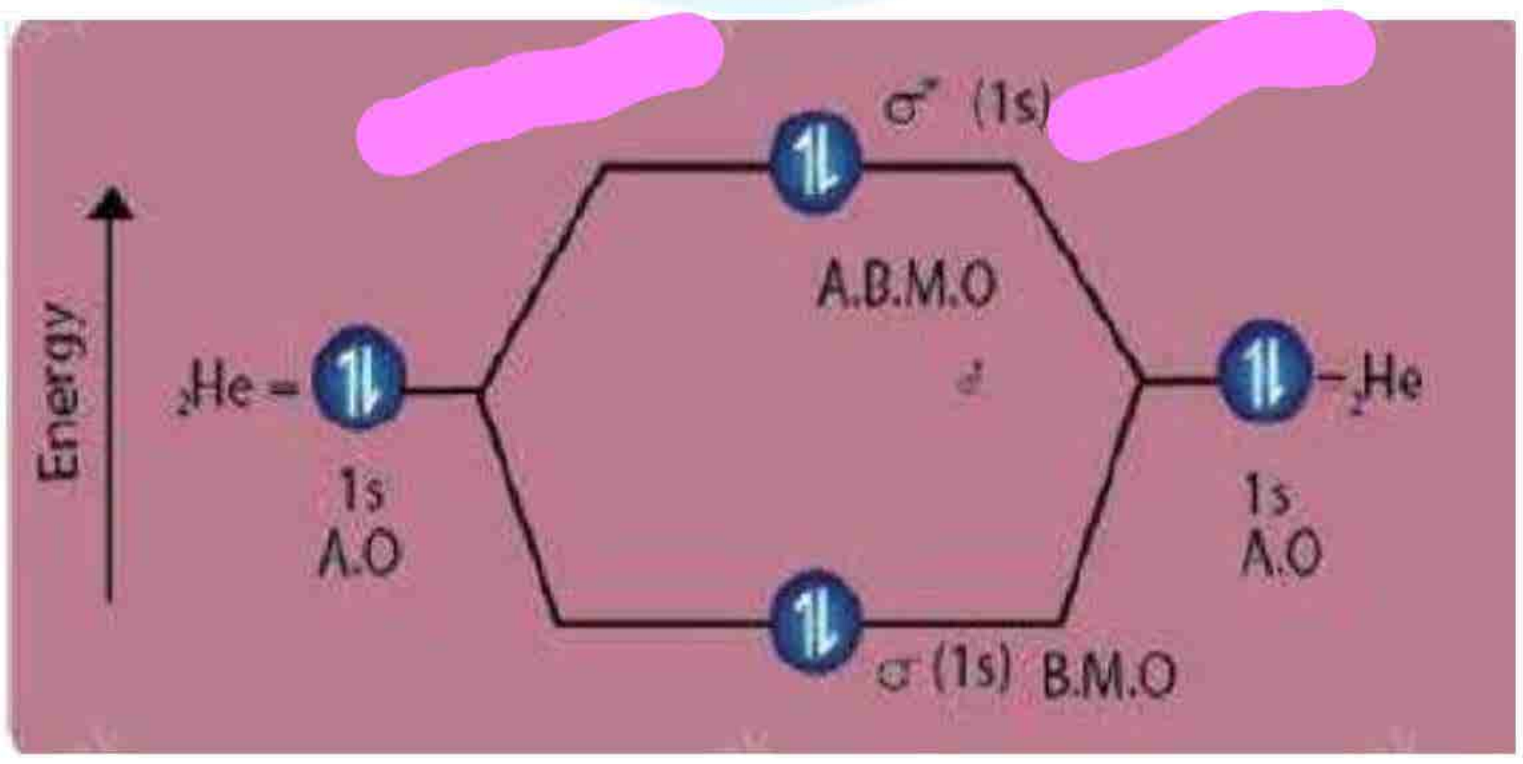


61. *Prove that no bond is formed in helium.*

Ans: The electronic configuration of He is $1s^2$. The 1s orbitals of He-atoms combine to form one bonding σ (1s) and one anti-bonding σ* (1s) orbitals.

Each He-atom contributes two electrons. Two electrons enter bonding molecular orbital σ(ls) and the remaining two go to antibonding σ* (Is) molecular orbital. The bond order for He_2 is zero i.e.

$$\frac{2-2}{2} = 0$$



62. *Explain structure of nitrogen according to MOT.*

Ans: Electronic configuration of N_2 molecule is:

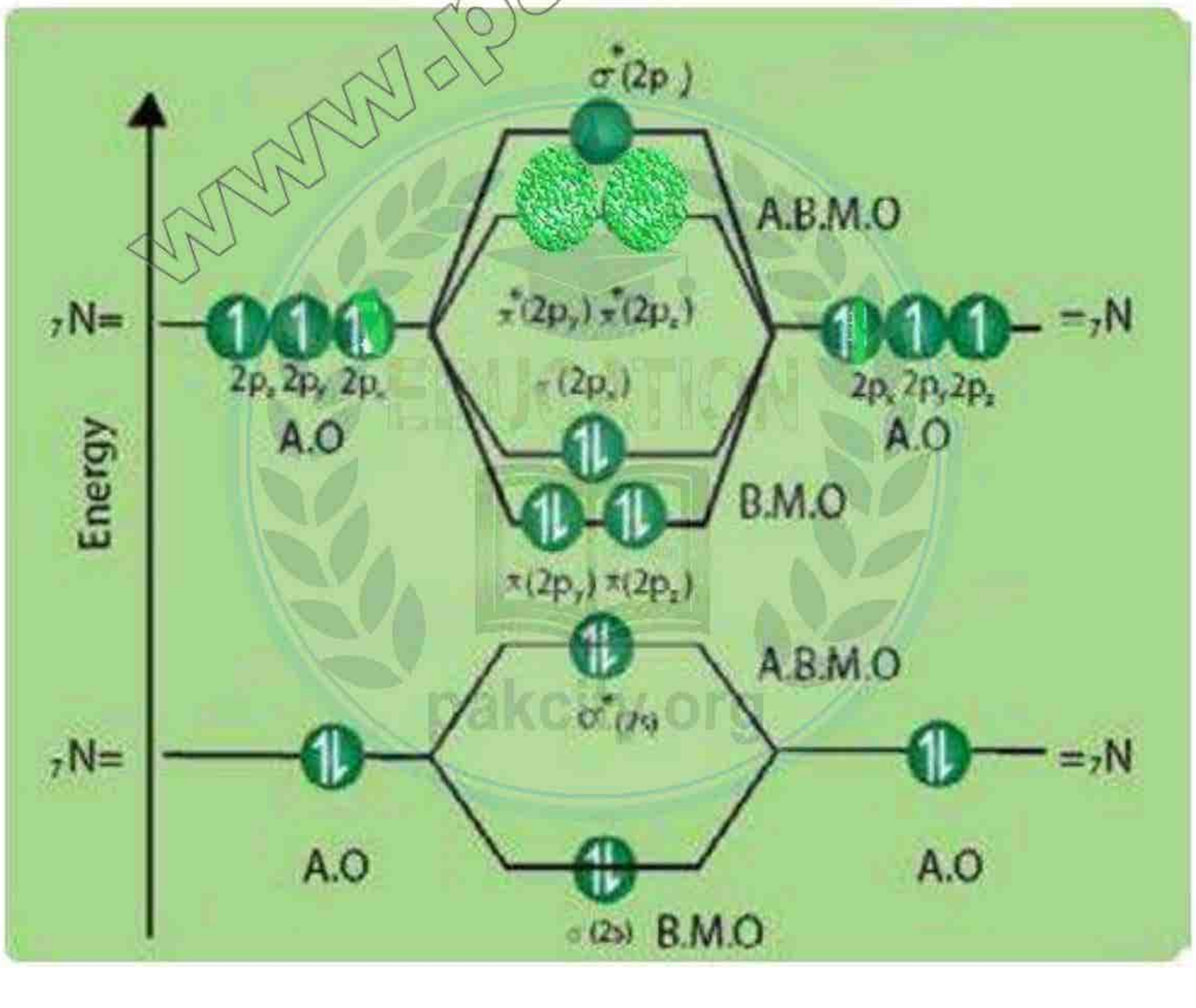
$$\sigma(1s^2)<\sigma^*(1s^2)<\sigma(2s^2)<\sigma^*(2s^2)<\pi(2p^2{}_y)=\pi(2p^2{}_z)<\sigma(2p^2{}_x)<\pi^*(2p_y)=\pi^*(2p_z)<\sigma^*(2p_x)$$

Six electrons enter into three outermostbonding orbitals while no electrons enter into anti-bonding orbitals.

Thus, the bond order in N_2 molecule is

$$\frac{6-0}{2}=\frac{6}{2}=3$$

which corresponds to the triple bond consisting of one sigma and two pi bonds. The bonddissociation energy of N_2 is very high, i.e. 941kJmol^{-1}.



63. *Explain structure of oxygen according to MOT.*

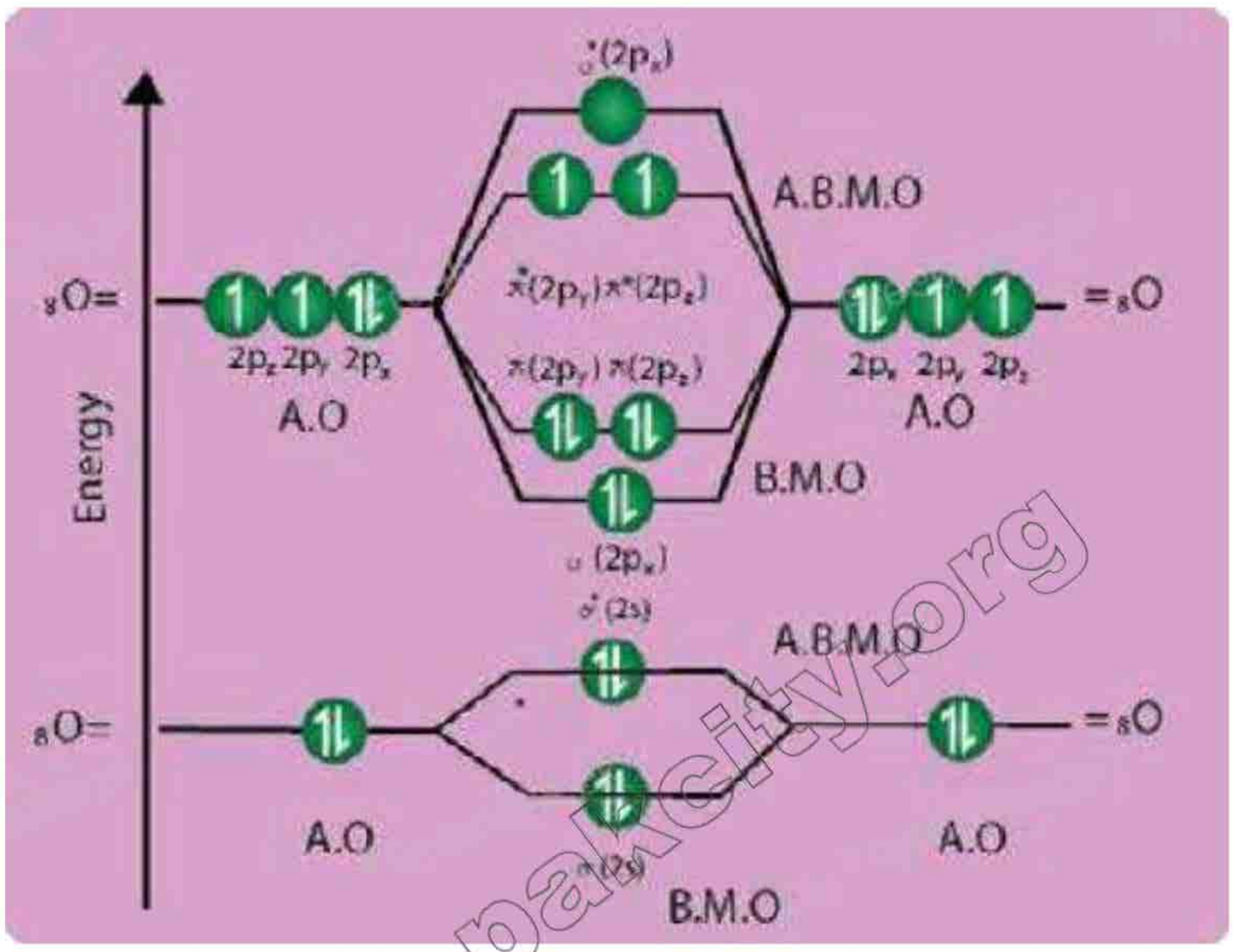
Ans: The electronic configuration of O_2 is:

$$\sigma(1s)^2<\sigma^*(1s)^2<\sigma(2s)^2<\sigma^*(2s)^2<\sigma(2px)<\pi(2py)^2=\pi(2pz)^2<\pi^*(2py)^1=\pi^*(2pz)^1<\sigma^*2px$$

The bond order of O_2 is:

$$\frac{6-2}{2}=2$$

This corresponds to a double bond. There are two unpaired electrons in anti-bonding molecular orbitals showing paramagnetic properties of oxygen. Liquid O_2 is attracted towards the magnet.



64. Differentiate between hybrid orbital and molecular orbital?

Ans:

Hybrid orbital	Molecular orbital
Hybrid orbitals are formed by the interactions of atomic orbitals in the same atom.	Molecular orbitals are formed by the interactions of atomic orbitals of two different atoms.
They have same energies	They may have different energies

65. Why the energy of anti-bonding molecular orbital is higher thancorresponding bonding molecular orbital?

Ans: Anti-bonding orbitals are higher in energy because there is less electron density between the two nuclei. An anti-bonding orbital is formed when two atoms approach each other and the overlap between atomic orbitals results in destructive interference. Thus, when the electrons in an anti-bonding orbital spend less time between the two nuclei, they are at a higher energy level.

66. Differentiate between atomic orbital and molecular orbital.

Ans:

Atomic orbital	Molecular orbital
Atomic orbital is the region, having the highest probability of finding an electron in an atom.	Molecular orbital is the region, having the highest probability of finding an electron in a molecule.
Formed by the electronic cloud around the atom.	Formed by the fusion of atomic orbitals that have nearly the same energy.

67. Define Dipole moment and give its S.I units.

Ans: The product of electric charge (q) and the distance between the positive and negative centers (r) is called dipole moment.

$$\mu = q \times r$$

The common unit of dipole moment is Debye (D). The S.I units of dipole moment is coulomb meter (Cm). 1D = 3.336×10^{-30} Cm

68. Define bond length. Give an example.

Ans: The distance between the nuclei of two atoms forming a covalent bond is called the bond length. For example, Cl-Cl bond length is 198 pm.

69. Define bond energy. Give an example.

Ans: The bond energy is the average amount of energy required to break all bonds of a particular type in one mole of the substance. For example, C-C bond energy is 348 kJ/mol.

70. Why the abnormality of bond length and bond strength in HI is less prominent than that of HCl.

Ans: Electronegativity of Cl is more than that of iodine that is why the difference of electronegativity between the HCl is more than of Hl between the bonded atoms. The decrease in polarity from the HCl to Hl indicates the increase in trend of equal sharing of electrons due to decreasing electronegativity between the bonded atoms Therefore, the bond length and bond strength of HCl is more prominent in its abnormality than that of Hl.

71. Why the dipole moment of CO_2 is Zero but that of SO_2 is 1.61 D?

Ans: The dipole moment of CO_2 is zero, as it has a linear structure where the dipoles being equal and opposite cancel out each other's effect.

SO_2 is also a triatomic molecule but it has a lone pair of electron due to which it has angular structure. Polarity is not cancelled out so it has a dipole moment of 1.61 D.

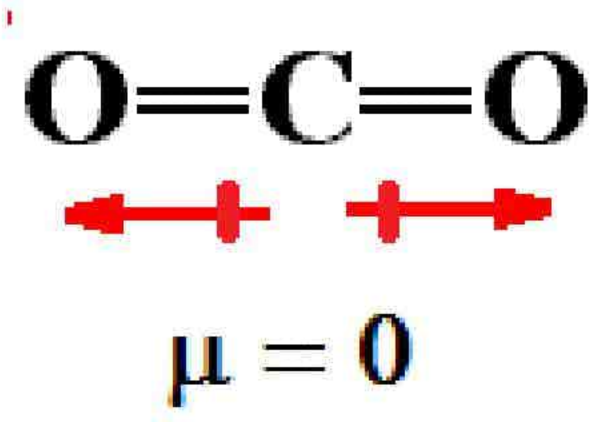
72. Why the dipole moment of SO_2 is 1.61D but that of SO_3 is zero?

Ans: SO_2 is a triatomic molecule but it has a lone pair of electrons, due to which it has angular structure. Polarity is not cancelled out, so it has a dipole moment of 1.61 D.

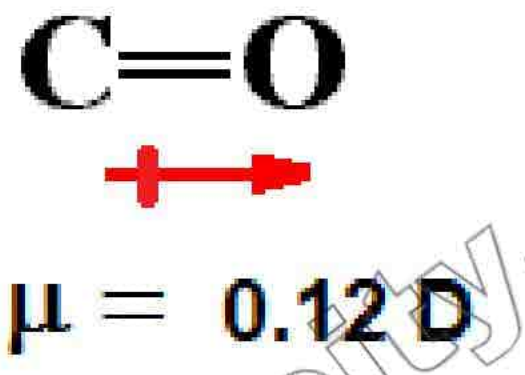
SO_3 has a dipole moment of 0. SO_3, a trigonal molecule, has no dipole moment because the bond dipoles cancel each other. This is because the negative and positive centers are both located in the same plane and therefore there is no permanent dipole therefore the molecule is non-polar.

73. Why the dipole moment of CO_2 is Zero but that of CO is 0.12 D?

Ans: The dipole moment of CO_2 is zero, as it has a linear structure where the dipoles being equal and opposite, cancel out each other's effect.



CO has a permanent pole, hence it shows a dipole moment of 0.12 D.



74. Why BF_3 is non-polar but SO_2 polar?

Ans: BF_3 has a symmetrical triangular planar molecule, having zero dipole moment so it is a non-polar molecule. On the other hand, SO_2 is also a triatomic molecule but it has a lone pair of electrons, due to which it has angular structure. Polarity is not cancelled out, so it has a dipole moment of 1.61 D and is polar in nature.

75. How the percentage of ionic character of covalent bond is determined by Dipole moment?

Ans: From the experimentally determined dipole moments, the percentage ionic character in a bond can be calculated.

$$\%\text{age ionic character} = \mu_{observed}/\mu_{ionic} \times 100$$

Example

The observed dipole moment of HF is 1.90D. The distance between the charges is 0.917×10^{-10} m. (unit positive charge = 1.6022×10^{-19} C)

$$\mu_{ionic} = q \times r$$

$$= 1.6022 \times 10^{-19}\ C \times 0.917 \times 10^{-10}\ m$$

$$= 1.469 \times 10^{-29}\ Cm = 4.4\ D \qquad (1D = 3.336 \times 10^{-30}\ Cm)$$

$$\%\text{age ionic character} = \mu_{observed}/\mu_{ionic} \times 100$$

$$= 1.90D/4.4D \times 100$$

$$= 43.2\%$$

76. Why the melting points, boiling points, heats of sublimation and heats of vaporization of electrovalent compounds are higher as compared with those of covalent compounds?

Ans: The melting points, boiling points, heats of sublimation and heats of vaporization of electrovalent compounds are usually higher as compared to covalent compounds. In ionic compound, strong inter-ionic forces are present. Large amount of energy is required to break these forces. Ionic compounds are solid in nature while covalent compounds are usually found in solids, liquids and gases. Covalent compounds have less attractive forces between molecules and have less melting points, boiling points, heats of sublimation and heats of vaporization. For example, melting point of an ionic compound NaCl is 801°C while melting point of a covalent compound H_2O is 0 °C.

77. Why it is impossible for CH_4 to make a coordinate covalent bond with H^+ ion while water and ammonia can do so?

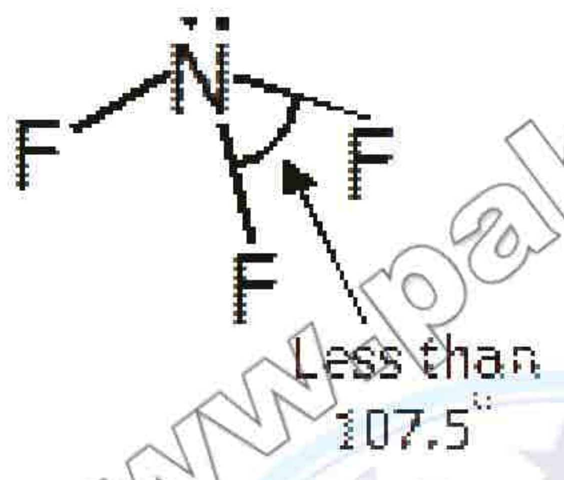
Ans: A coordinate covalent bond is formed between two atoms, when the shared pair of electrons is donated by one of the bonded atoms, called donor and accepted by the other bonded atom, called the acceptor.

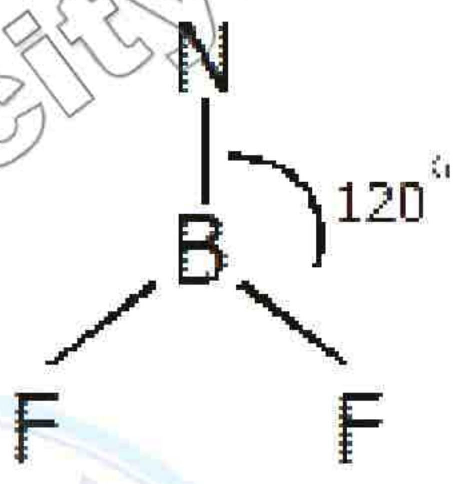
H^+ ion acts as an electron pair acceptor. Ammonia and water both have lone pair of electrons, which they can donate to H^+ ion to form coordinate covalent bond.

Methane CH_4 has only bond pairs, and no lone pair is available so it cannot form coordinate covalent bond.

78. The molecules NF_3 and BF_3 all have molecular formula of the type XF_3. But they have different structural formulas. Keeping in view VSEPR theory, sketch the shape of each molecule and explain the origin of differing in shapes.

Ans: The difference in their shapes or geometries is due to a lone pair which is present on nitrogen in NF_3 but not on boron in BF_3. This lone pair repels the three N-F bonds in NF_3 downwards and shape becomes trigonal pyramidal, whereas, the shape of BF_3 remains triangular planar because of absence of such repulsions.





79. Why the lone pairs of electrons occupy more space than bond pairs?

Ans: A lone pair is attracted by only one nucleus, on the other hand, a bonding electron pair is attracted by both nuclei of atoms. Because a lone pair experiences less nuclear attraction, its electronic charge is spread out more in space than that of bond pair. As a result, electron pairs are more influential and exert greater repulsive forces on bond pairs, thus tend to compress the bond pairs occupying more space.

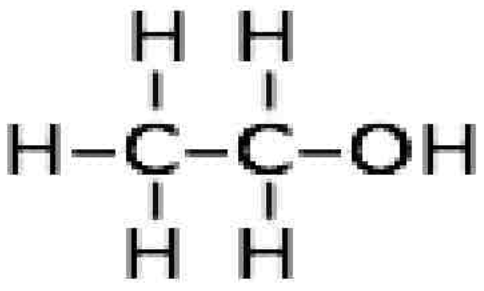
80. On what factors strength of bond depends?

Ans: The strength of bond depends on

i. Electronegativity
ii. Size of atom
iii. Bond length

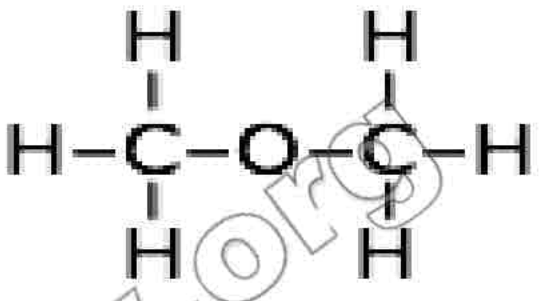
81. Why ionic compounds do not show the phenomenon of isomerism, but covalent compounds do?

Ans: Since ionic bond is non-directional, ionic compounds do not exhibit the phenomenon of isomerism. Covalent bonds are rigid and directional as overlapping of orbitals occurs along definite axes. This leads to the possibility of a variety of isomerism. For example, structural isomerism is shown by the compound, C_2H_6O.



ethanol

C_2H_6O



dimethyl ether

C_2H_6O

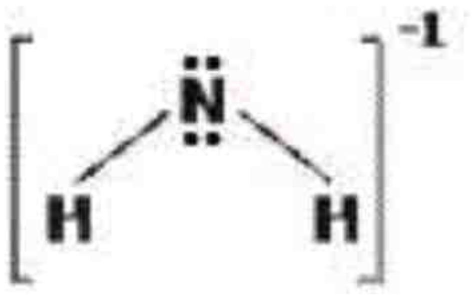
82. How the type of bonding affects the solubility of compounds.

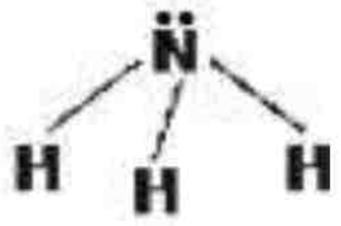
Ans: Mostly ionic compounds are soluble in water but insoluble in non-aqueous solvents. When a crystal of ionic substance is placed in water, the polar water molecules detach the cation and anion from the crystal lattice by their electrostatic attraction and salt is dissolved in water. If the attraction of the water molecules cannot overcome the attraction between the ions then dissolution does not occur. For the same reasons, non-polar solvents like benzene and hexane do not dissolve ionic compounds. Covalent compounds dissolve easily in non-polar organic solvents like benzene, ether etc. Here the attractive forces of solvent molecules are enough for overcoming the intermolecular forces of attraction.

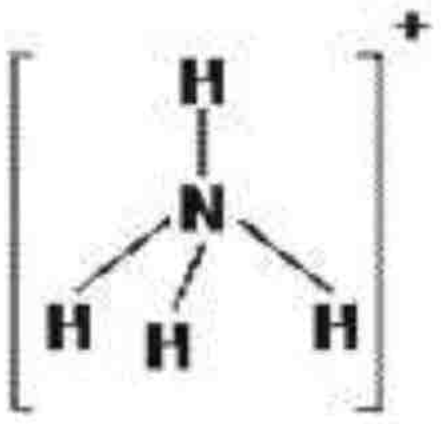
Compounds like glucose, sucrose, urea etc. dissolve in water due to hydrogen bonding.

83. The species NH_2^- NH_3, NH_4^+, have bond angles of 105°, 107.5° and109.5° respectively. Justify these values by drawing their structures.

Ans: This is due to presence of lone pairs on central atom of these species (except NH_4^+). They repel the bond pairs, thereby, decreasing the bond angles. Perfect tetrahedral bond angle is shown by NH_4^+ only.







84. The linear geometry of $BeCl_2$ suggests that central Be atom is sp- hybridized. What type of hybridization a central atom undergoes, when the atoms bonded to it are located at the corners of (a) an equilateral triangle (b) a regular tetrahedron and (c) triangular bipyramidal?

Ans: **(a)**

In case of equilateral triangle, the central atom should be sp^2 hybridized e.g. Boron in BF_3.

(b)

In case of regular tetrahedron, the central atom should be sp^3 hybridized e.g. Carbon in CH_4.

(c)

In case of Triangular bipyramidal, the central atom should be dsp^3 hybridized e.g.Phosphorous in PCl_5.

85. How does molecular orbital theory explain the paramagnetic characterof O_2, O_2^{2+} and O_2^{2-} species?

Ans:

O_2:

O_2 will be paramagnetic because it has 2 unpaired electrons in anti-bonding molecular orbitals.

O_2^{2+}:

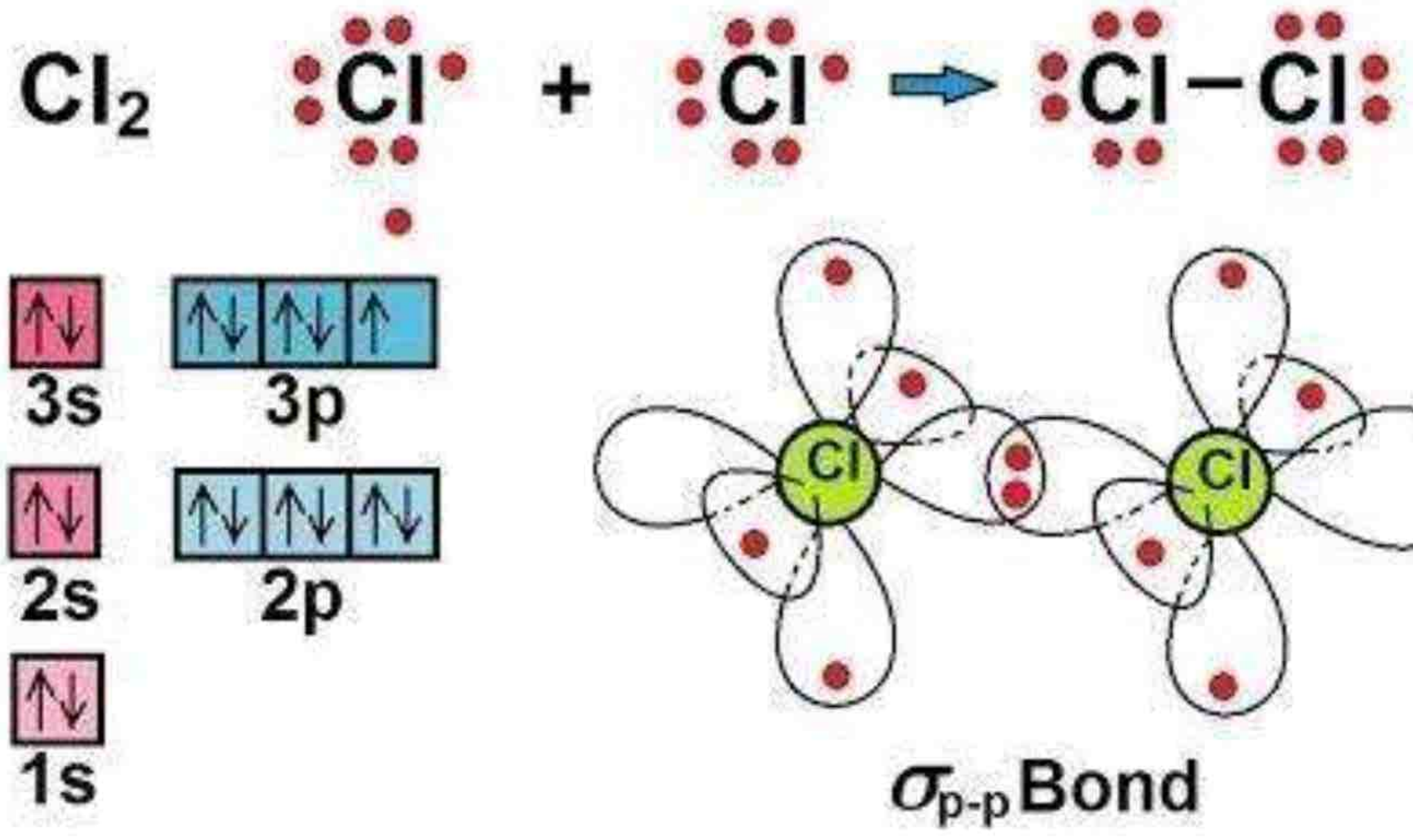
O_2^{2+} will be diamagnetic because it has no unpaired electrons. O_2 after losing two electrons becomes O_2^{2+}

O_2^{2-}:

O_2^{2-} will be again diamagnetic because it has no unpaired electrons. O_2 after gaining two electrons becomes O_2^{2-}.

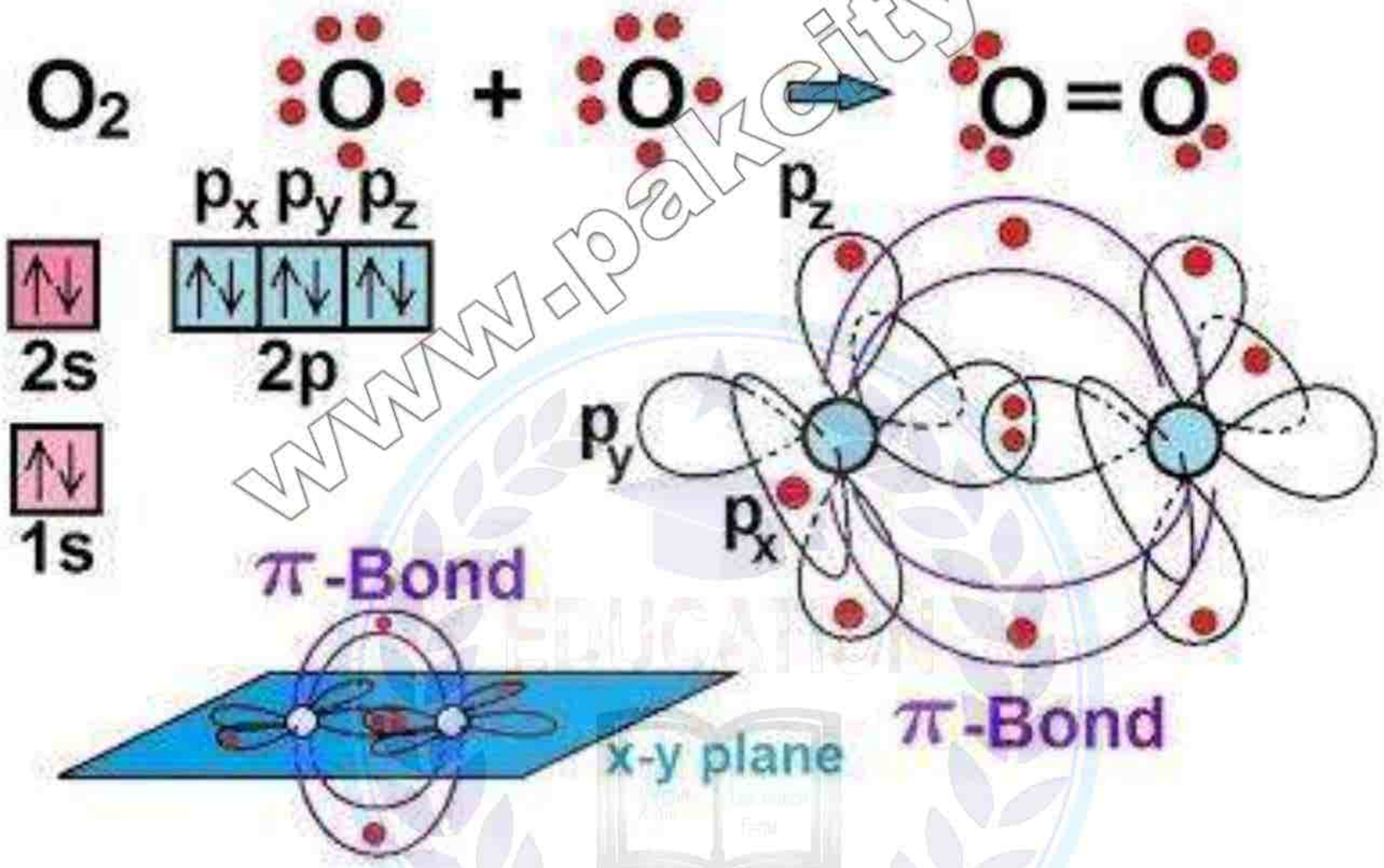
86. How the bonding in Cl_2 is explained according to valence bond theory?

Ans: In Cl_2 formation, two 3px orbitals of two chlorine atoms overlap with each other to make a single covalent bond.



87. How bonding in O_2 is explained according to valence bond theory?

Ans: In O_2 formation, two 2px orbitals of two oxygen atoms overlap with each other to make a single sigma covalent bond and later on two 2py orbitals and two 2pz orbitals engage in the formation of two pi bonds.

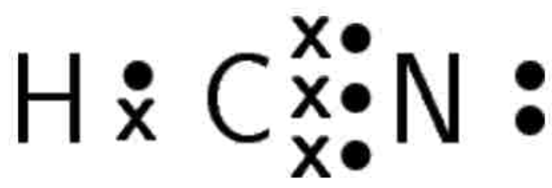


88. Write the Lewis structures for the following compounds:

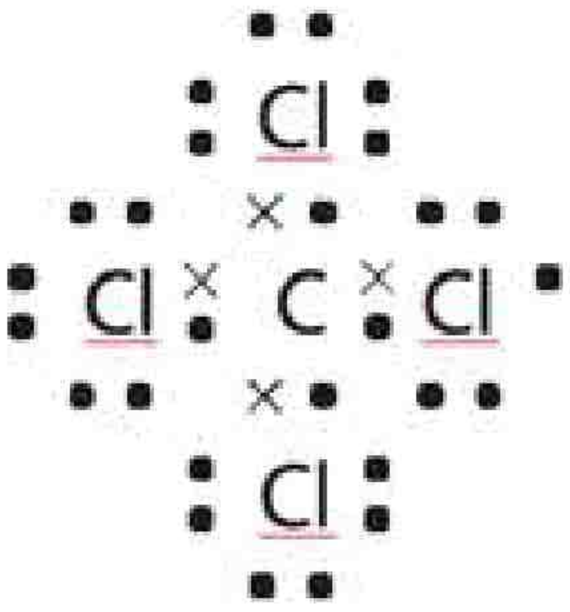
(i) HCN (ii) CCl_4 (iii) CS_2 (iv) $H_3N \rightarrow AlF_3$ (v) NH_4OH (vi) H_2SO_4 (vii) H_3PO_4 (viii) $K_2Cr_2O_7$ (ix) N_2O_5 (x) $Ag(NH_3)_2NO_3$

Ans:

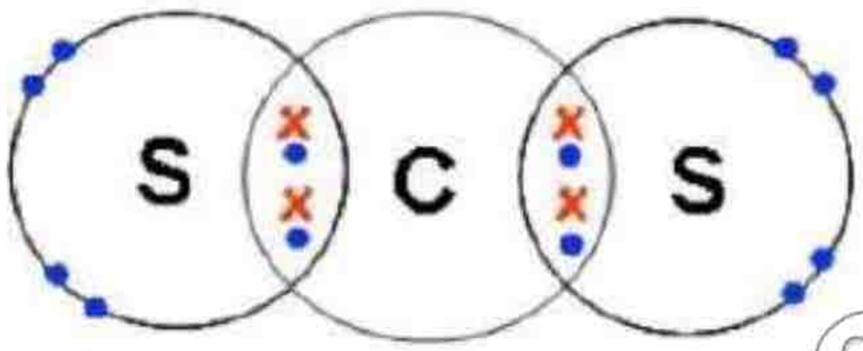
HCN



CCl_4

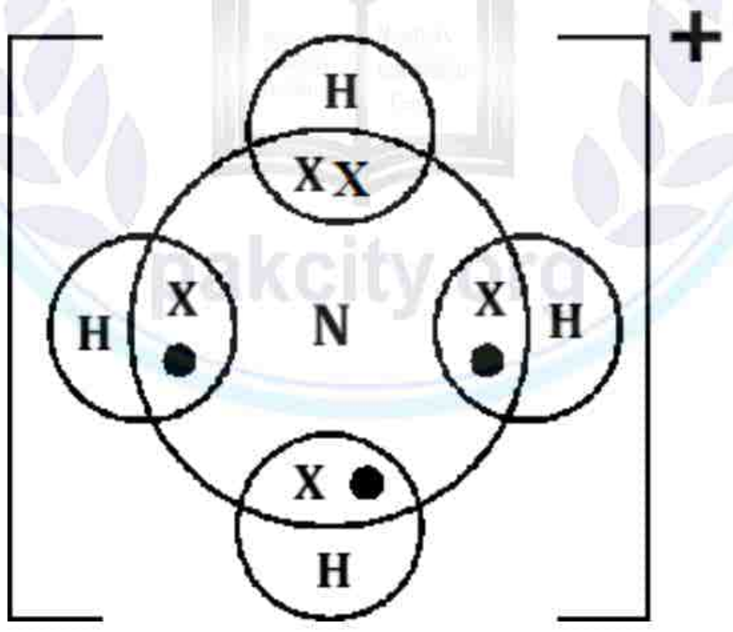


CS_2

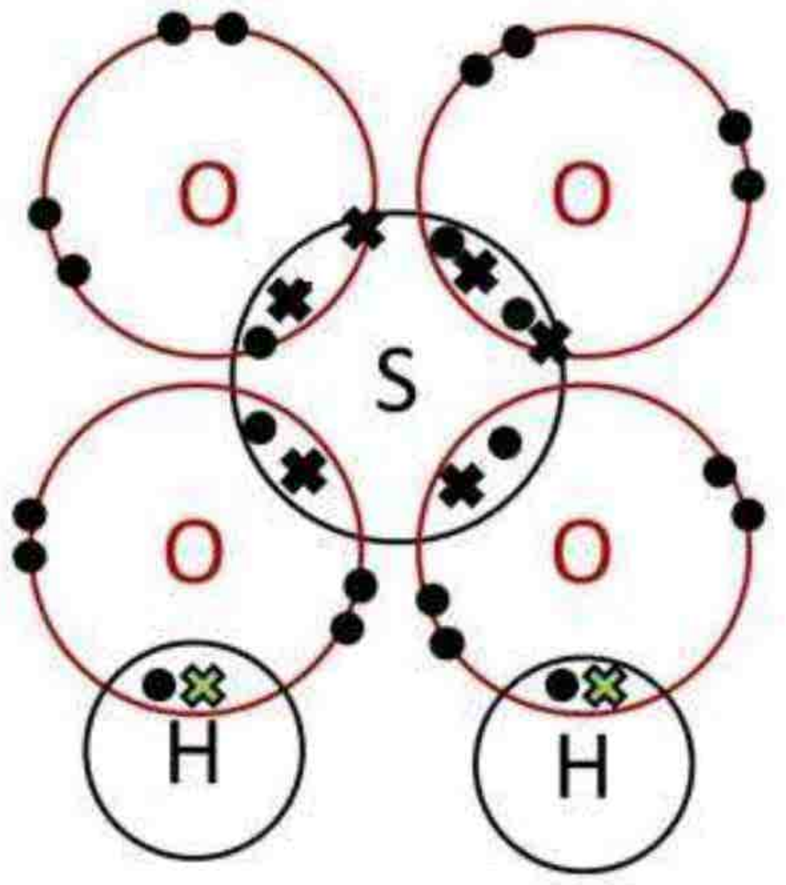


$H_3N \rightarrow AlF_3$

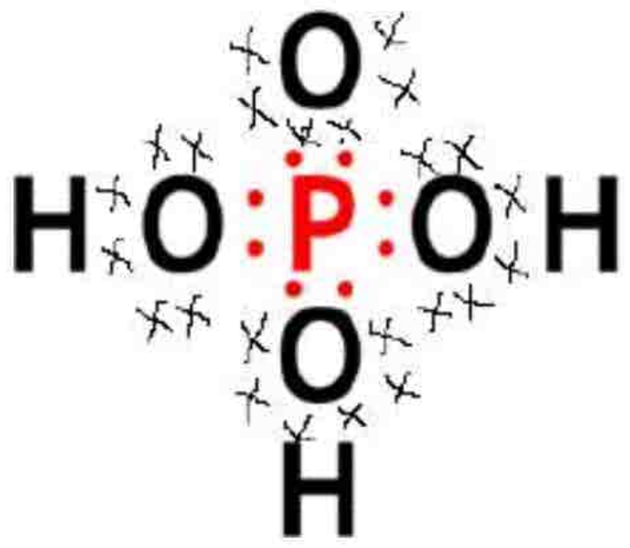
NH_4OH



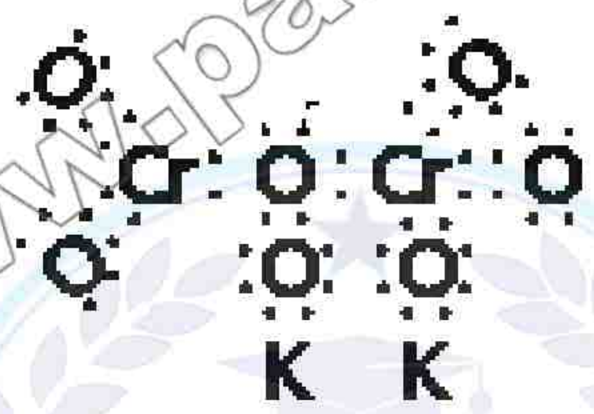
H_2SO_4



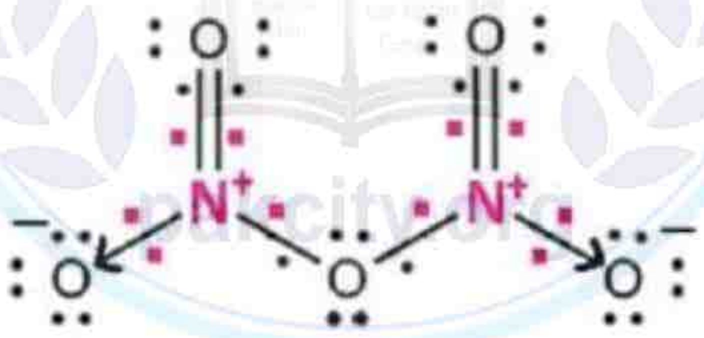
H_3PO_4



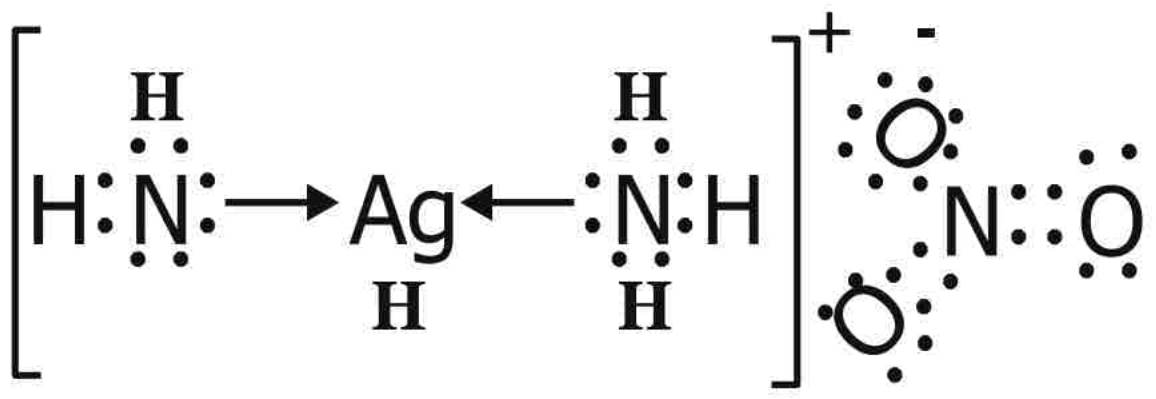
$K_2Cr_2O_7$



N_2O_5



$Ag(NH_3)_2NO_3$



89. How do you compare the bond strength of:

(i) Polar and non-polar molecules (ii) σ and π bonds?

Ans: **(i)**

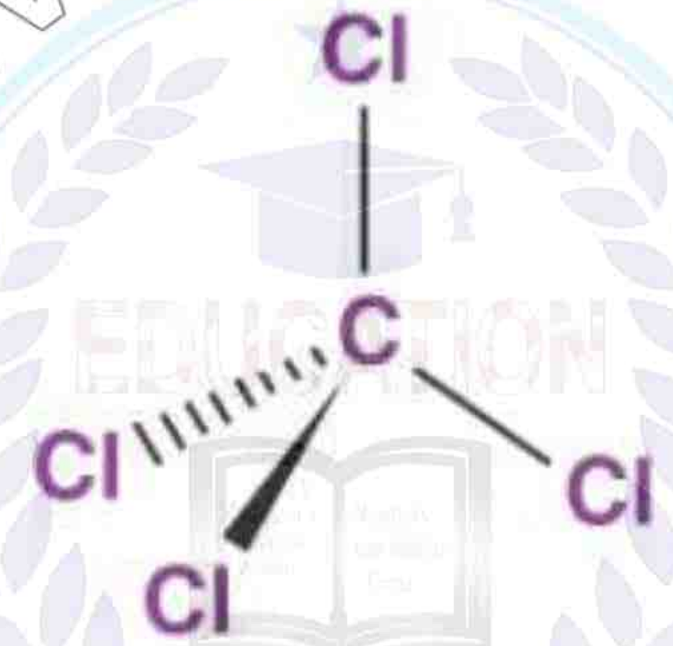
The polar bond has more strength as compared with non-polar bond because of attractions between partial positive and negative ends of the molecule.

(ii)

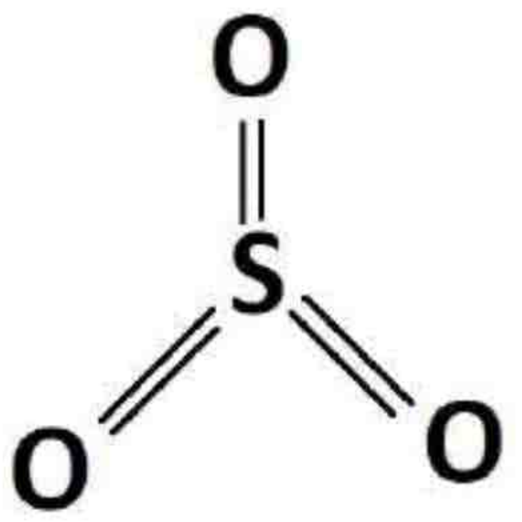
Sigma bond is formed by the linear overlap of half-filled orbitals of atoms, thus, it is less diffused (less spread) and stronger than pi bond which is formed by parallel overlap of half-filled orbitals above and below the plane of molecule making it more diffused and weak.

90. Which of the following molecules will be polar or non-polar, sketch the structures and justify your answer? (i) CCl_4 (ii) SO_3 (iii) NF_3 (iv) SO_2

Ans: **CCl_4**: non-polar due to zero dipole moment resulting in tetrahedral structure.

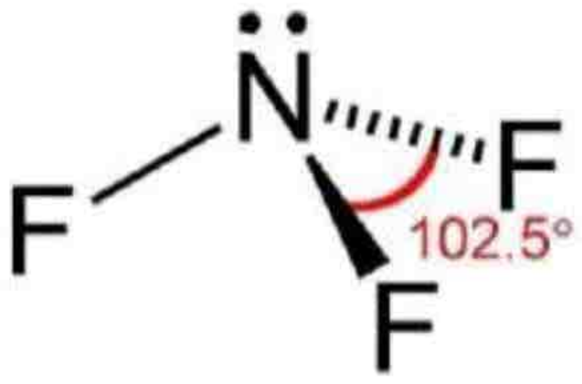


SO_3: non-polar due to zero dipole moment resulting in triangular planar structure.

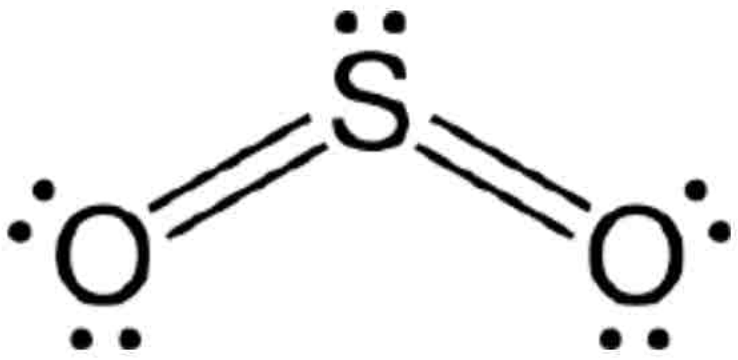


Sulphur Trioxide

NF_3 : polar due to dipole moment resulting in trigonal pyramidal structure.



SO_2 shall be polar due to resultant dipole moment (1.61D) of its angular structure.



91. How Pauling calculated inter-ionic radii?

Ans: Let us consider, the positive and negative ions, which are held together by electrostatic forces of attraction in a crystal lattice.

r^+ and r^- are the values of radii of cation and anion, respectively. The interionic distance 'R' will be:

$$R = r^+ + r^-$$

Pauling was able to determine the distance between K^+ and Cl^- ions in potassium chloride crystal and found that it was equal to the sum of the radii of the two ions.

$$R = 133\text{pm} + 181\text{ pm} = 314\text{ pm}$$

Thus, the ionic radius appeared to be an additive property. Pauling extended this concept to other K^+ salts and calculated the radii of other ions from the relationship:

$$r^- = R - r^+$$

Similarly, the ionic radii of different cations can also be determined.

92. How the covalent radius of bonded atoms can be calculated?

Ans: The covalent radius of an atom can be used to determine the covalent radius of another atom. For example, the experimentally determined bond length of C-Cl in CH_3CI is 176.7 pm. The covalent radius of Cl-atom is 99.4 pm. Covalent radius of C-atom can be calculated by subtracting this value from C-Cl bond length. So, the covalent radius of C-atom = 176.7- 99.4 = 77.3 pm.

93. The dipole moments of CO_2 and CS_2 are zero but that of SO_2 is 1.61D.

Ans: The dipole moments of CO_2 and CS_2 are zero because of their linear structures in which resultant dipole moment is zero due to cancellation of dipole moment on both sides. In SO_2, a lone pair is present on sulphur introducing repulsions on bond pairs and molecular geometry becomes angular which will have a resultant dipole moment of 1.61 D.