

Q1. Give Lewis structure of SO_4^{2-} .

Q2. Write a neutral molecule which is isoelectronic with ClO^- .

Q3. Write Lewis symbols for the following atoms and ions:
S and S^{2-} ; Al and Al^{3+} ; H and H^- .

Q4. How many types of bonds are present in NH_4Cl ?

Q5. What is the total number of electrons in NO_3^- ?

Q6. Write electron dot structures of CO and AlCl_3 .

Q7. Although very useful in a large number of cases, the octet rule has many exceptions. Give two examples to support this statement.

Q8. How many σ and π -bonds are present in $\text{CH}_2 = \text{CH} - \text{CH} = \text{CH}_2$?

Q9. Why is NaCl harder than sodium metal?

Q10. Why is solid NaCl non-conductor of electricity?

Q11. Explain the formation of a chemical bond.

Q12. Write the Lewis dot structures of BeF_2 and SiCl_4 .

Q13. Why does BF_3 behave as Lewis acid?

Q14. Indicate the number of σ and π bonds in the molecule $\text{CH}_2 = \text{C} = \text{CH}_2$.

Q15. Calculation of formal charge on O atoms of O_3

$$\begin{array}{c} 2 & 1 & 3 \\ \vdots & \ddot{\vdots} & \vdots \\ \text{:O} & \text{:O} & \text{:O:} \\ \vdots & \ddot{\vdots} & \vdots \end{array}$$

Q16. Give electron dot structure for carbon suboxide, C_2O_2 . Write its structural formula.

Q17. Give Lewis structure of SOCl_2 .

Q18. Give Lewis structure of ClO_4^- .

Q19. Calculation of formal charge on Cl atom in HClO_4

$$\begin{array}{c} \text{:O:} \\ \text{H} & \text{:O:Cl:O:} \\ \text{:O:} \end{array}$$

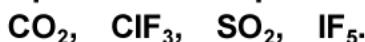
Q20. Calculation of formal charge on P atoms in H_3PO_4

$$\begin{array}{c} \text{H:O:} \\ \text{H} & \text{:O:P:O:H} \\ \text{:O:} \end{array}$$

Q21. Which of the following do not obey Octate rule:

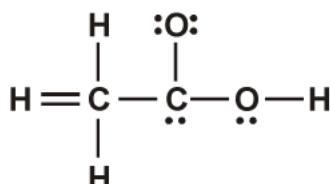


Q22. Which of the following are superoctate compound:



Q23. Write the significance of a plus and a minus sign shown in representing the orbitals.

Q24. The skeletal structure of CH_3COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.



Q25. Which end of ICI will be positive and which will be negative and why? Is it covalent or ionic?

Q26. Give two differences between σ and π -bonds.

Q27. Using Lewis dot symbol show electron transfer between cation and anion of following pairs.

Q28. Draw Lewis structure of following element and ions.

(a) P and P^{3-} (b) C and C^+ (c) P and P^+ (d) H and H^-

Q29. Three elements have the following Lewis symbols:



- (a) Write the group of the element in the periodic table.
- (b) Which elements are expected to form ions? What is the expected charge on the ions?

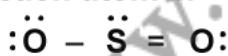
Q30. Use the concept of formal charges to determine the most stable Lewis dot structure for CO_2 in the following.

(a) $O = C = O$ (b) $O = O - C$ (c) $O - C = O$
(d) $C - O \equiv O$ (e) $O - O = C$

031. Calculate the formal charges of the atoms in the nitrate ion (NO_3^-).

Q32. Calculate the formal charges of the atoms in the carbonate ion (CO_3^{2-}).

Q33. Calculate the formal charge on each atom in



Q34. Define octet rule. Write its significance and limitations.

Q35. Explain the formation of H_2 molecule on the basis of valence bond theory.

Q36. What do you understand by bond pairs and lone pair electrons? Illustrate by giving one example of each type.

Q37. What is the total number of sigma and pi bonds in the following molecules?

(a) C_2H_2 (b) C_2H_4

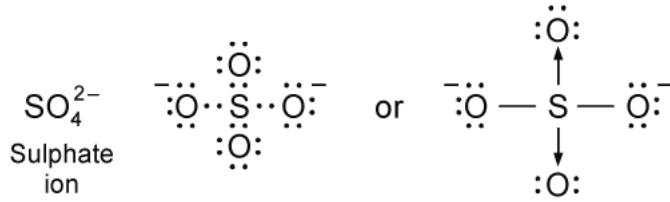
Q38. Draw diagrams showing the formation of a double bond and a triple bond between carbon atoms in C_2H_4 and C_2H_2 molecules.

Q39. Considering x-axis as the internuclear axis which out of the following will not form a sigma bond and why?

(a) 1s and 1s (b) 1s and $2p_x$ (c) $2p_y$ and $2p_z$ (d) 1s and 2s

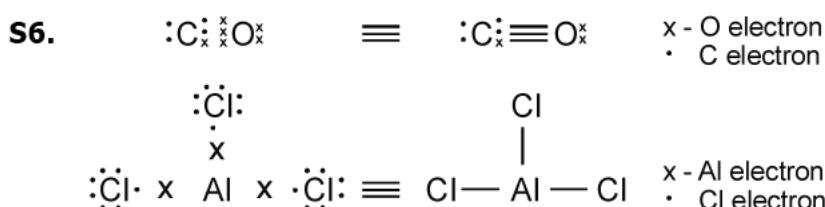
Q40. Use Lewis symbols to show electron transfer between the following atoms to form cations and anions: (a) K and S (b) Ca and O (c) Al and N.

Q41. Write the favourable factors for the formation of ionic bond.

S1.

S2. XeO , ClF .

Species	^{16}S	^{13}Al	^1H
Atom	$\ddot{\text{S}}$	$\ddot{\text{Al}}\cdot$	$\ddot{\text{H}}$
Ion	$[\ddot{\text{S}}]^{2-}$	$[\text{Al}]^{3+}$	$[\ddot{\text{H}}]^-$

S4. Three types bonds (three covalent, one coordination, one ionic bond.)

S5. 32 electrons.

S7. In AlCl_3 : Al share 6 electrons.

 In PCl_5 : P share 10 electrons.

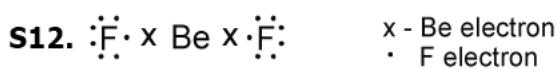
S8. σ bond = 9; π bond = 2.

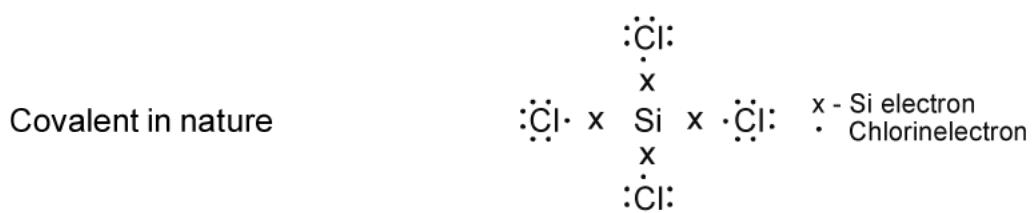
S9. NaCl is ionic in nature and ions (Na^+ Cl^-) has much high lattice energy due to electrostatic force. while Na metal has respectively less lattice energy between atoms.

S10. Because NaCl is ionic in nature having strong lattice bond. Hence ions are not mobile in nature, and cannot transfer electrons (electrons are not mobile in nature) to conduct electricity.

S11. Chemical bonds are two in types.

- By transformation of electrons and cation anion formed which have electrostatic force-called ionic bond
- By sharing of electrons between two atoms. Two electrons of two atoms share their orbitals by decreasing their energies in this case net attractive and repulsive force is equal, called covalent bond





S13. Because in BF_3 , Boron has empty p -orbital. i.e., Boron has 6 electrons after bonding three F -atoms.

S14. σ bond = 6; π bond = 2.

S15. Formal charge on $\text{O}^1 = 6 - \frac{1}{2} (6) - 2 = +1$.

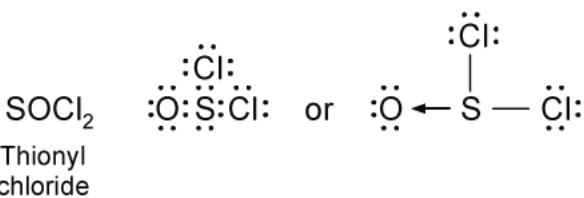
Formal charge on $\text{O}^2 = 6 - \frac{1}{2} (4) - 4 = 0$.

Formal charge on $\text{O}^3 = 6 - \frac{1}{2} (2) - 6 = -1$.

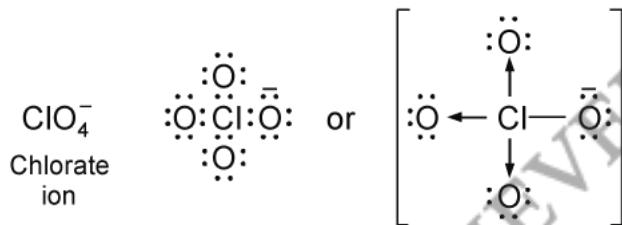
S16. $\text{O}=\text{C}=\text{C}=\text{C}=\text{O}$



S17.



S18.



S19. Formal charge on Cl = $7 - \frac{1}{2} (8) - 0 = 3$.

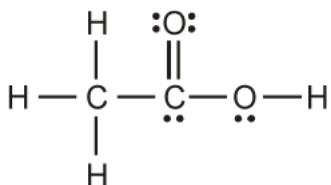
S20. Formal charge on P = $5 - \frac{1}{2} (8) - 0 = 1$.

S21. SF_4 , SO_2 , SO_3 , NO , BF_3 .

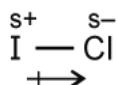
S22. ClF_3 , SO_2 , IF_5 .

S23. An orbital is a pictorial representation of wave function. The plus and minus signs do not represent charges but instead they represent + and -ve nature of the wave function ' ψ '.

S24. The skeletal arrangement of the atoms in the above structure is correct, however, electronic arrangements is not in accordance with Lewis concept. The correct structure of acetic acid is



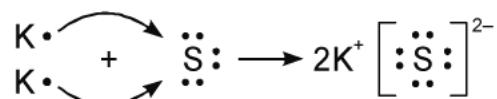
S25. ICl is covalent compound, which is polar in nature and I is respectively electropositive the Cl hence I is positive end while Cl is negative end



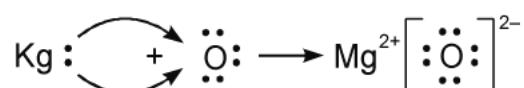
S26.	σ bonds	π bonds
i.	In σ bonds orbitals mix. end to end of two different atoms	In π bonds orbitals mix. side by side of two different atoms.
ii.	σ bonds determined the shape of molecule	π bonds has not contribution to determine shape

S27. (a) K and S

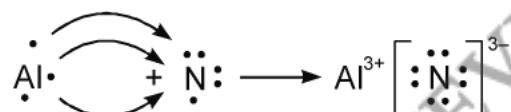
Potassium is monovalent while S is bivalent.



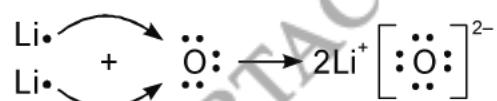
(b) Mg and O both bivalent



(c) Al and N. both trivalent



(d) Li is monovalent O is divalent



S28. (a) $\ddot{\text{:P:}}$ and $\left[\ddot{\text{:P:}}\right]^{3-}$ (b) $\ddot{\text{C}}$ and $\ddot{\text{C}}^+$

(c) $\ddot{\text{:P}}$ and $\left[\ddot{\text{:P}}\right]^+$

(d) $\text{H}\cdot$ and $[\ddot{\text{H}}]$

S29. (a) Number of dots in the Lewis symbol is equal to the number of valence electrons in an atom and thus we find that:

(i) In $\text{X}\cdot$, the number of dot is one and number of valence electron is also one. Therefore, X is an element of group 1 of the periodic table.

(ii) In $\text{Y}\cdot$, there are four valence electrons. Therefore, Y is an element of group 14 of the periodic table.

(iii) In $\text{Z}\cdot$, there are seven valence electrons. Therefore, Z is an element of group 17 of the periodic table.

(b) (i) $\text{X}\cdot$ will form a cation by losing one electron.



(ii) $\text{Z}\cdot$ will form an anion by gaining one electron.

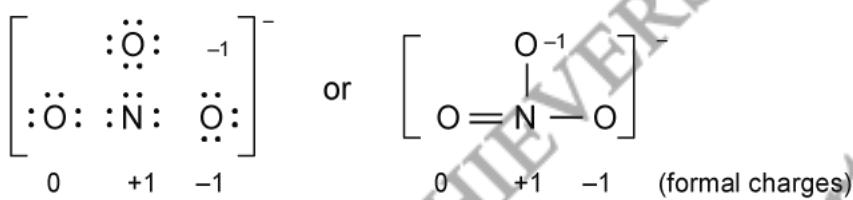


S30. The formal charges for each atom in each structure are given as follows.



It is seen that the formal charges are minimized when the carbon atom is in the middle and forms double bonds to each oxygen. Thus structure (a) represents a stable Lewis structure of CO_2 .

S31. Lewis dot structure of NO_3^- ion is written as



$$\text{Formal charge of an atom} = N_v - N_{ub} - \frac{1}{2} N_b$$

(a) Formal charge of N atom $= 5 - 0 - \frac{1}{2}(8) = 5 - 4 = +1$

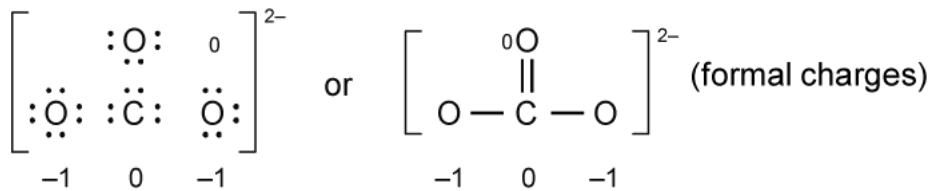
(b) Formal charge of O atom (left to N) $= 6 - 4 - \frac{1}{2}(4) = 6 - 6 = 0$

(c) Formal charge of O atom (right to N) $= 6 - 6 - \frac{1}{2}(2) = 6 - 7 = -1$

(d) Formal charge of O atom (above N) = $6 - 6 - \frac{1}{2}(2) = 6 - 7 = -1$.

Sum of the formal charges = $0 + 1 - 1 - 1 = -1$

S32. Lewis dot structure of CO_3^{2-} ion is written as:



Formal charge of an atom = $N_v - N_{ub} - \frac{1}{2}N_b$

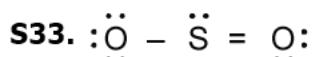
(a) Formal charge of C atom = $4 - 0 - \frac{1}{2}(8) = 4 - 4 = 0$

(b) Formal charge of O atom (left to C) = $6 - 6 - \frac{1}{2}(2) = 6 - 6 - 1 = -1$

(c) Formal charge of O atom (right to C) = $6 - 6 - \frac{1}{2}(2) = 6 - 6 - 1 = -1$

(d) Formal charge of O atom (above C) = $6 - 4 - \frac{1}{2}(4) = 6 - 4 - 2 = 0$

Sum of the formal charges = $-1 + 0 - 1 + 0 = -2$



Formal charge on atom

$(\text{Nr}) - (\text{Nnb}) - \frac{1}{2}(\text{Nb})$ and both atoms.

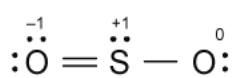
= No. of valence electrons - No. of non-bonding electrons - $\frac{1}{2}$ No. of bonding electrons.

FC on S atom = $6 - 2 - \frac{1}{2}(6) = +1$

FC on double bonded O atom = $6 - 4 - \frac{1}{2}(4) = 0$

FC on single bonded O atom = $6 - 6 - \frac{1}{2}(2) = -1$

So, we may write :



Note: Single bonding O form co-ordination bond, and its own electron does not take part in bonding.

S34. Octet rule: According to octet rule atoms take part in bonding (covalent or ionic) to complete 8 electrons in outermost shell to get nearest noble gas stable-configuration.

Limitation of Octet rule:

- Some atoms form stable covalent compounds in which their octet is incomplete, less than 8-electrons in outermost shell called octet deficient compounds. e.g., LiCl, Li-has 4 electrons.
- Some atoms form stable covalent compounds in which they have more than 8-electrons in outermost shell called super octet compounds. e.g., PCl₅, P-has 10 electrons in outermost shell.

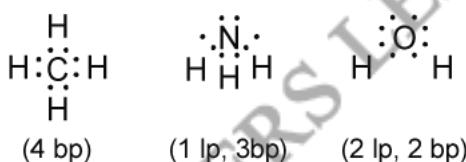
S35. Hydrogen atom has one electron which is present in its 1s-orbital.

Assume two atom H_A and H_B with their corresponding electrons e_A and e_B respectively, when two atoms start to approaching the electron starts coming under to influence of the nucleus of the other. i.e., their orbitals begin to interact new electrostatic force is created.

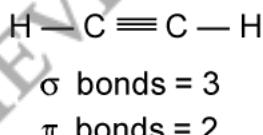
Attractive force between electron e_A and nucleus of H_B, electron e_B and nucleus of H_A repulsive force between electron e_A and e_B and nucleus H_A and H_B.

In bond formation magnitude of attractive forces is greater than the repulsive forces.

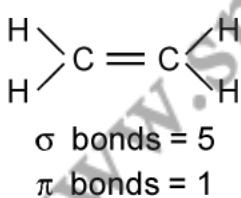
S36. The pair of electrons which are shared by the two atom to constitute a covalent bond is called **bond pair**. On the other hands the electron pair of the valence shell which is not used in sharing/bonding is called **lone pair**. For example, in CH₄ there are only four bond pairs, in NH₃ molecule, there are 3 bond pairs, and one lone pair and in water molecule there are two lone pairs and 2 bond pairs as shown below:



S37. In C₂H₂



In C₂H₄



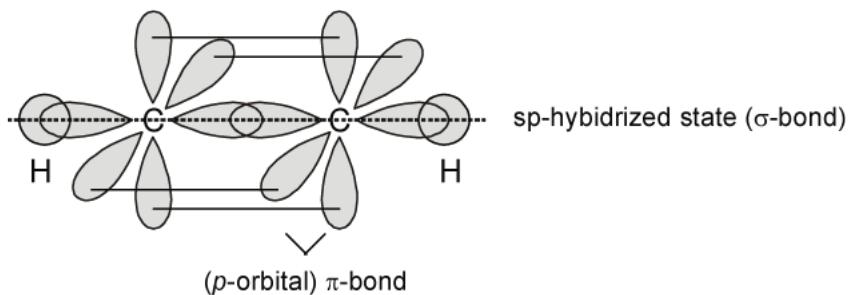
S38. Formation of C ≡ C in ethyne:

In C₂H₂ C is *sp*-hybridized

₆C = 1s² 2s² 2px¹ 2py¹ 2pz⁰ Ground state

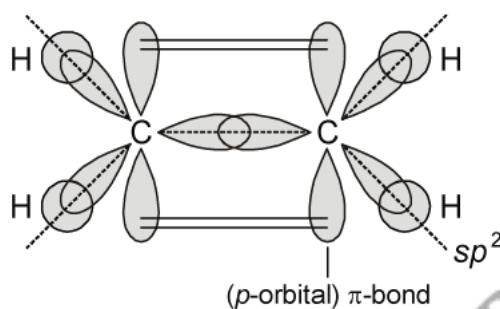
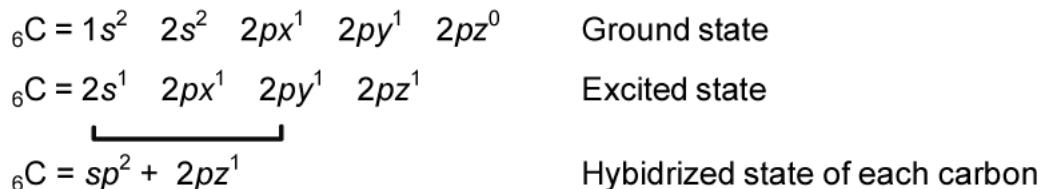
₆C = 2s¹ 2px¹ 2py¹ 2pz¹ Excited state

$\begin{array}{c} \boxed{1} \\ \text{6} \end{array}$ C = *sp* + 2py¹ 2pz¹ Hybridized state of each carbon



Formation of C = C in ethene:

In C_2H_4 each carbon is sp^2 -hybridized



S39. The sigma bond is formed by axial overlap and it can be formed by following orbitals.

$2p_x$ and $2p_z$ cannot form σ bond because they are not in axis in x-axis.

S40. (a) $2\dot{K} + \ddot{S} \rightarrow (K^+)_2 \left[\ddot{S}\right]^{2-}$

(b) $\text{Ca} + \ddot{\text{O}} \rightarrow \text{Ca}^{2+} [\ddot{\text{O}}]^{2-}$

(c) $\cdot\ddot{\text{Al}}\cdot + :\ddot{\text{N}}: \longrightarrow \text{Al}^{3+} [\ddot{\text{N}}:]^{3-}$

S41. Ionic bond is formed by transference of electrons from one atom to another. The favourable conditions for its formation are:

- (i) Low ionisation enthalpy of element forming cation.
- (ii) More negative value of electron gain enthalpy of element forming the anion and
- (iii) High value of lattice enthalpy of the compound formed.

Q1. Compare the H — N — H bond angle in the following molecules and ions:

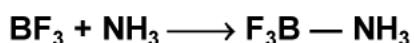


Q2. The H — S — H bond angle in H_2S is 92.2° whereas the H — O — H bond angle in H_2O is 104.5° , why?

Q3. Which of the following has maximum bond angle?

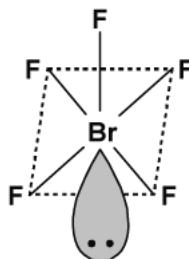


Q4. Is there any changes in the hybridisation of B and N atoms as a result of the following reaction:

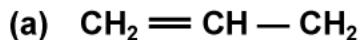


Q5. Arrange H_2O , NH_3 and CH_4 in the decreasing order of bond angle.

Q6. Explain the shape of BrF_5 .



Q7. What is hybrid state of each carbon in



Q8. What is hybrid state of central atom in the following?



Q9. What can you say about the shapes of SO_2 and NF_3 molecules from the information that both have net dipole moment?

Q10. Give the structure of an anion which is isostructural with BF_3 .

Q11. What is the state of hybridisation of carbon in CO_3^{2-} ion?

Q12. Give shape and hybridization of SF_6 .

Q13. Give shape and hybridization of SF_4 .

Q14. Give shape and hybridization of SCl_2 .

Q15. Give shape of PCl_3 and hybridization.

Q16. Give shape of SO_4^{2-} and hybridization.

Q17. Give shape of CO_3^{2-} and hybridization

Q18. Give shape of $\dot{\text{N}}\text{H}_3$ and hybridization

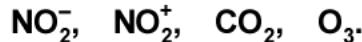
Q19. Give shpae and hybidrization of NH_4^+ .

Q20. Give shpae and hybidrization of BeF_3^- .

Q21. What type of hybridisation explains the trigonal bipyramidal shape of SF_4 ?

Q22. What is the type of hybrid orbitals associated with B atom in BH_4^- ?

Q23. Which of the species have similar shape and why?



Q24. Give hybridization and shape of ICl_2^\ominus .

Q25. Give hybridization and shape of IF_7 .

Q26. Give hybridization and shape of IF_5 .

Q27. Give hybridization and shape of ClF_3 .

Q28. Give hybridization and shape of XeF_2 .

Q29. Give hybridization and shape of XeF_6 .

Q30. Give hybridization and shape of XeOF_4 .

Q31. Give hybridization and shape of ICl_4^\ominus .

Q32. Give hybridization and shape of XeF_4 .

Q33. BeF_2 molecule is linear while SF_2 is angular though both are triatomic.

Q34. Why NF_3 is pyramidal while BF_3 is triangular planar, though both are tetra atomic molecules?

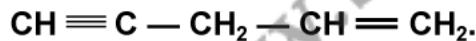
Q35. What is the total number of sigma and pi bond in the following molecules:



Q36. (a) How many sigma and pi bonds are there in the ollowing molecule.



(b) Which type of hybrid orbitals are used by the second carbon atom in the following molecule.



Q37. Bond angle in NH_3 is more than in H_2O . Why?

Q38. Why is BeCl_2 linear whereas SnCl_2 angular molecule?

Q39. Explain why PCl_5 is trigonal bipyramidal whereas IF_5 is square pyramidal?

Q40. Explain the structure of C_2H_2 with orbital diagram.

Q41. Describe the hybridisaton in case of PCl_5 . Why are the axial bonds longer as compared to equatirial bonds?

Q42. Give the formula of the noble gas species which is isostructural with

(a) IBr_2^\ominus (b) ICl_4^\ominus (c) BrO_3^\ominus

Q43. Bond angle in PH_4^+ is higher than that in PH_3 . Why?

Q44. Explain hybridization of P in PX_5 in box notation and draw shape.

Q45. Draw the shapes of (a) POCl_3 ; (b) XeF_4 .

Q46. BCl_3 is planar but anhydrons AlCl_3 is tetrahedral.

Q47. How is VB theory different from Lewis concept in regard to the formation of covalent bond?

Q48. Arrange the following in order of decreasing bond angle around N atom, NO_2 , NO_2^+ , NO_2^- .

Q49. How do bond length of C — H bond differ in C_2H_6 , C_2H_4 and C_2H_2 ?

Q50. Why is that in SF_4 molecule, the lone pair of electrons occupy equatorial position in preference to axial position? What is the shape o the molecule?

Q51. What is the hybrid state of C and O atoms in CO_2 . Draw its orbital overlap diagram.

Q52. Give the hybridization and structures of the following species:

(a) NO_3^\ominus (b) ClO_4^\ominus

Q53. Apart from tetrahedral geometry, another possible geometry for CH_4 is square planar with four H atoms at the cormers of the square and the C atom at its centre. Explain why CH_4 is not square planar.

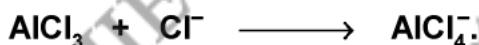
Q54. Which hybrid orbitals are used by carbon atoms in the following molecules?

(a) $\text{H}_3\text{C} — \text{CH} = \text{CH}_2$ (b) $\text{CH}_3 — \text{CH}_2\text{OH}$ (c) CH_3CHO (d) CH_3COOH

Q55. Is there any change in the hybridisation of B and N atoms as a result of the following reaction?



Q56. Describe the change in hybridisation (if any) of the Al atom in the following reaction.

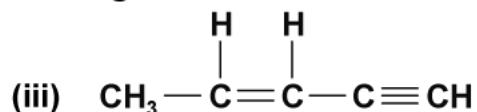


Q57. (a) Draw the Lewis structure of tetracyanoethylene and point out the total number of sigma and pi-pond.

(b) What is total number of sigma and pi-bonds in the following structures

(i) $\text{C}_2\text{H}_3\text{Cl}$

(ii) CH_2Cl_2



Q58. Consider the structure, $\overset{1}{\text{C}}\overset{2}{\text{H}}_3\overset{2}{\text{C}}\overset{3}{\text{H}}_2\overset{3}{\text{C}}\overset{4}{\text{O}}\overset{4}{\text{H}}_2\overset{5}{\text{C}}\overset{5}{\text{H}}\overset{6}{\text{C}}\overset{6}{\text{H}}$ and answer the following

(a) Calculate number of σ - and π -electrons in the structure.

(b) Which atoms in the structure have same hybrid state.

(c) Arrange the atoms C_2 , C_3 , C_5 in decreasing order of s-character of bonding orbitals.

Q59. Draw the probable structure of I_3^+ and I_3^- species and predict their shapes.

Q60. What is hybrid state of S in SO_2 molecule. Using the box representation describe the formation of SO_2 molecule on the basis of hybridisation. What type of bonds are present in the molecule.

Q61. Draw molecular structures of XeF_2 , XeO_2F_2 , XeF_4 .

Q62. Indicate the type of bonds present in NH_4NO_3 and state the mode of hybridisation of two N atoms in it.

Q63. ClF_3 exists while FCl_3 does not. Why?

Q64. The bond angle FNF in NF_3 is smaller than HNH angle in NH_3 but FPF bond angle in PF_3 is relatively larger than HPH angle in PH_3 . How can you account for it.

Q65. Arrange the following species in order of decreasing bond angles around the central atom in each set.

(a) CH_3 , NH_3 , H_2O , BCl_3 , C_2H_2

(b) NH_3 , $\text{N}\bar{\text{H}}_2$, NH_4^+

(c) H_2O , H_2Se , H_2Te , ZnCl , CO_2

S1. Bond angle: NH_4^+ > NH_3 > NH_2^-
(108) (107) (104)

S2. Because S is bigger in size than O and lonepair-bond, pair repulsion in H_2S is less than H_2O .

S3. CO_2 has maximum angle 180°



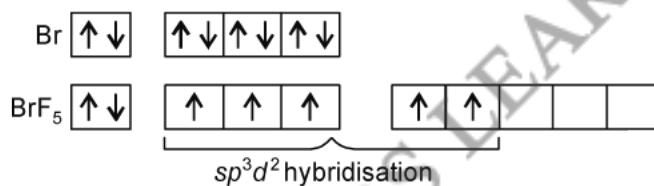
S4. In BF_3 , the hybridisation of B is sp^2 and in NH_3 , the hybridisation of N is sp^3 .

After the reaction, the hybridisation of B changes to sp^3 but the hybridisation of N remains unchanged (sp^3).

S5. Bond angle decreases as:

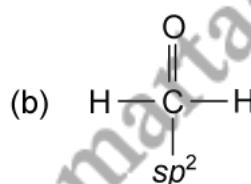
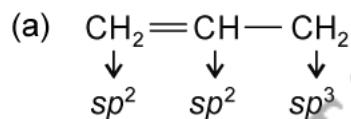


S6. BrF_5 involves sp^3d^2 Hybridisation and has one lone pair of electrons as:



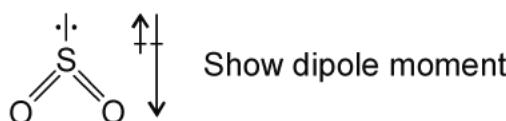
The shape is square pyramidal.

S7.

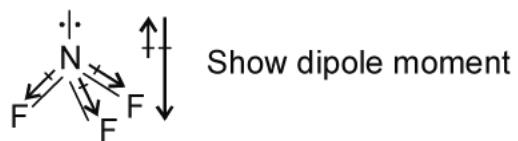


S8. NO_3^- — sp^2
 BF_4^- — sp^3
 PF_5 — sp^3d^2
 IF_7 — sp^3d^3

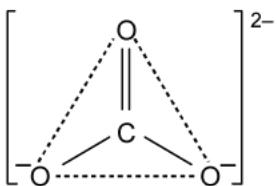
S9. SO_2 is bent in shape.



NF_3 is bent pyramidal in shape

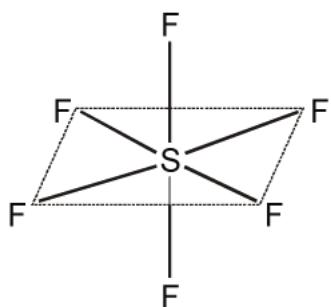


S10. CO_3^{2-}



S11. sp^2 .

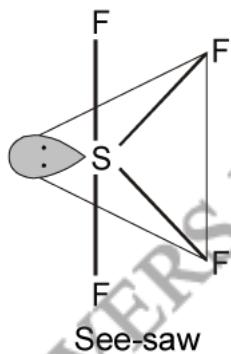
S12.



Hybridization – sp^3d^2

Shape – Octahedral

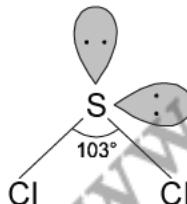
S13.



Hybridization – sp^3d

Shape – See-saw

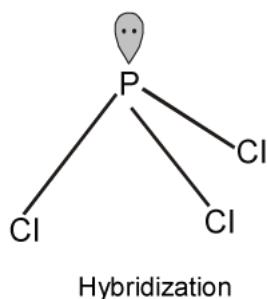
S14.



Hybridization – sp^3

Shape – Bent

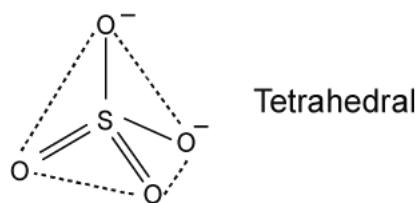
S15.



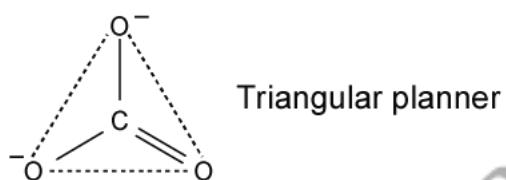
Hybridization – sp^3

Shape – bent pyramidal

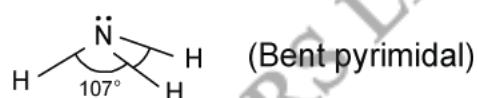
S16. SO_4^{2-} – sp^3 hybridization



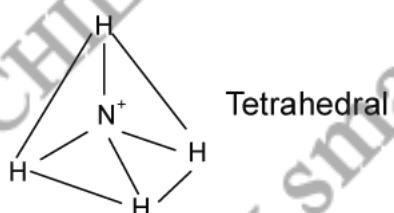
S17. CO_3^{2-} – sp^2 hybridization



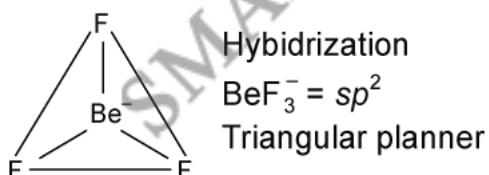
S18. NH_3 – sp^3 hybridization.



S19. NH_4^+ – Hybridization sp^3



S20.

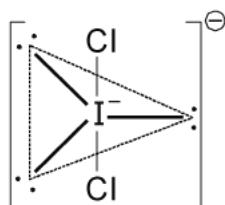


S21. sp^3d .

S22. sp^3 .

S23. NO_2^+ and CO_2 has linear shape, because of same hybridization and same number of electron on nitrogen and carbon.

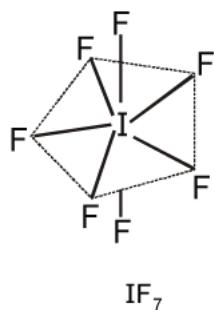
S24.



Hybridization – sp^3d

Shape – linear

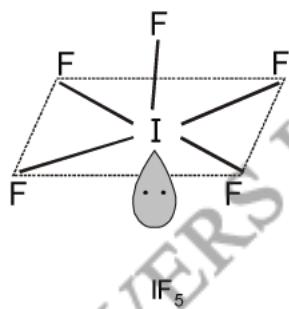
S25.



Hybridization – sp^3d^3

Shape – Pentagonal bipyramidal

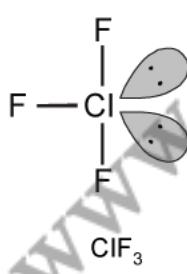
S26.



Hybridization – sp^3d^2

Shape – Square pyramidal

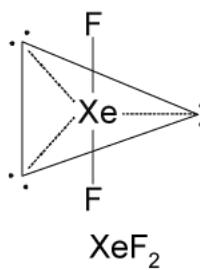
S27.



Hybridization – sp^3d

Shape – T-shaped

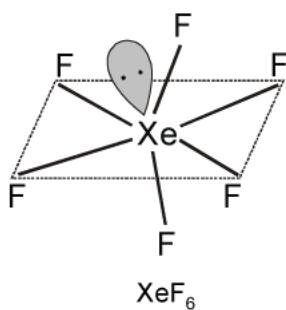
S28.



Hybridization – sp^3d

Shape – Linear

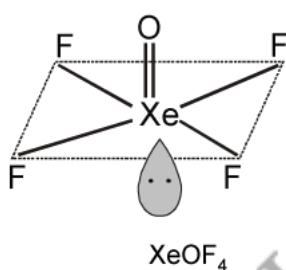
S29.



Hybridization – sp^3d^3

Shape – Distorted octahedral

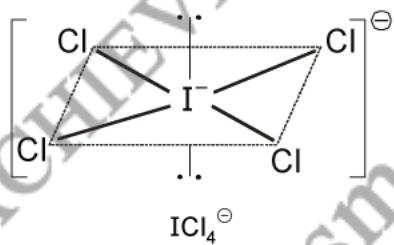
S30.



Hybridization – sp^3d^2

Shape – Square pyramidal

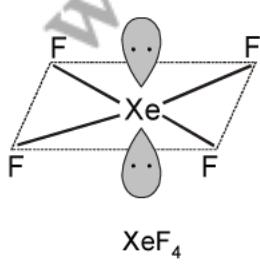
S31.



Hybridization – sp^3d^2

Shape – Square planar

S32.



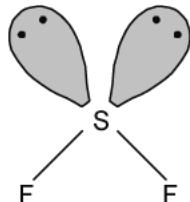
Hybridization – sp^3d^2

Shape – Square Planner

S33. In BeF_2 , Be is surrounded only by 2 electron bond pairs and therefore, the molecule is linear. On the other hand, in SF_2 , the central atom S, is surrounded by 2 bond pairs and 2 lone pairs. Since it is surrounded by 4 electron pairs, the geometry is expected to be tetrahedral but the larger repulsions of lone pair-bond pair, the bond angle is less than 109.5° and the resultant molecule is angular.

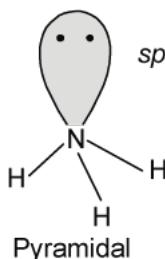


2 bond pairs
(Linear)



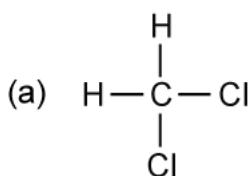
2 bond pairs, 2 lone pairs
(Angular)

S34. In NF_3 , nitrogen involves sp^3 hybridisation and one position is occupied by a lone pair. Therefore, the molecule is pyramidal. But in BF_3 , B involves sp^2 hybridisation having triangular planar geometry.

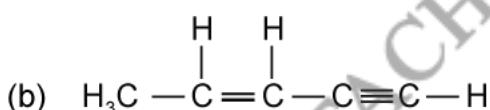


sp^3 hybridisation of N

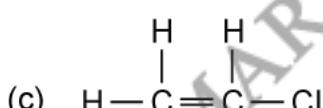
S35.



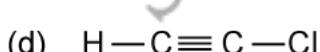
No. of σ bonds = 4
No. of π bonds = 0



No. of σ bonds = 10
No. of π bonds = 3



No. of σ bonds = 5
No. of π bonds = 1



No. of σ bonds = 3
No. of π bonds = 2

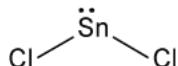
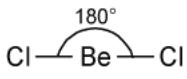
S36. (a) σ bonds = 10; π bonds = 3

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

S37. The difference in bond angles is due to the different numbers of lone pairs and bond pairs in the two species. In NH_3 , the N atom has two lone pairs and three bond pairs while in H_2O , the O atom has two lone pairs and two bond pairs. The repulsive interactions of lone pairs and bond pairs in water are relatively more than those in NH_3 . Hence, bond angle around central atom in water is relatively smaller than that in NH_3 molecule.

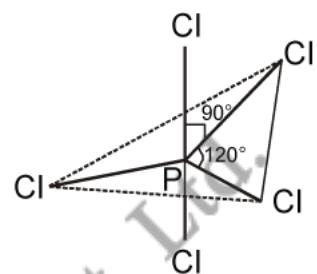
S38. Because BeCl_2 has sp hybridization and Be has not lone pair.

while SnCl_2 is sp^2 hybridized and Sn has one lone pair. Due to lone pair-bond pair repulsion SnCl_2 is bent

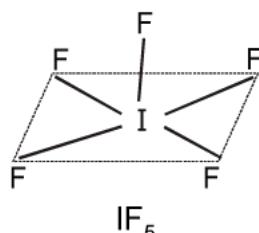


S39. PCl_5 has sp^3 hybridization and having bond-pair – bond pair repulsion.

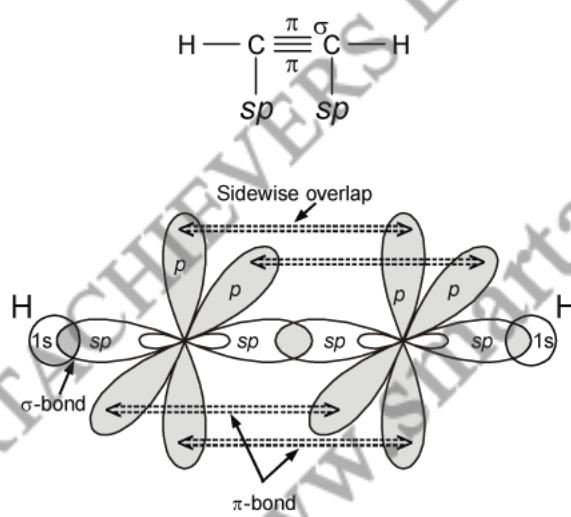
Three chlorine atoms arrange in equatorial position and form triangular shape (120°) while two chlorine atoms arrange axial position (90°) i.e., triangular bipyramidal shape is formed.



IF_5 has sp^3d^2 hybridization. I has one lone-pair five bond pair. Five bond (—F) arrange in square planner. One bond (—F) arrange in primidal (axil). One lone pair arrange itself in second pyrimid. Which has no. repulsion i.e., square pyamidal shape is formed.

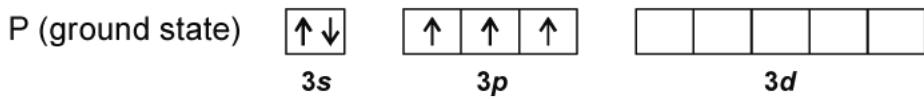


S40. In C_2H_2 each carbon is sp hybridized and P_y and P_z form two π bonds.



S41. In the formation of PCl_5 , phosphorus atom assumes sp^3d hybrid state. The longer nature of axial bonds is due to relative stronger repulsive interactions experienced by the axial bond pairs as compared to equatorial bond pairs.

Formation of PCl_5 :

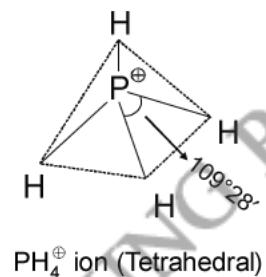
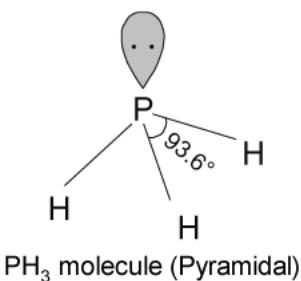




S42. Isostructural noble gas molecules:

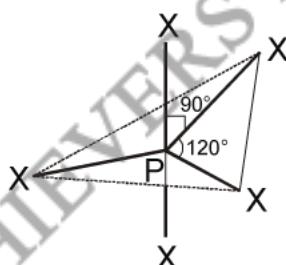
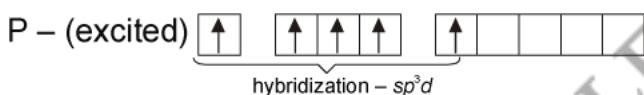
(a) IBr_2^\ominus : XeF_2
 (b) ICl_4^\ominus : XeF_4
 (c) BrO_3^\ominus : XeF_3

S43. P in PH_3 is sp^3 -hybridised. It has three bond pairs and one lone pair around P. Due to stronger lone pair-bond pair repulsions than bond pair-bond pair repulsions, the tetrahedral angle decreases from $109^\circ 28'$ to 93.6° . As a result, PH_3 is pyramidal. However, when it reacts with a proton, it forms PH_4^\oplus ion which has four bond pairs and there are no lone pair-bond pair repulsions. PH_4^\oplus , therefore assumes tetrahedral geometry with a bond angle of $109^\circ 28'$. Hence bond angle of PH_4^\oplus is higher than in PH_3 .



S44.

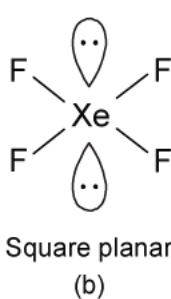
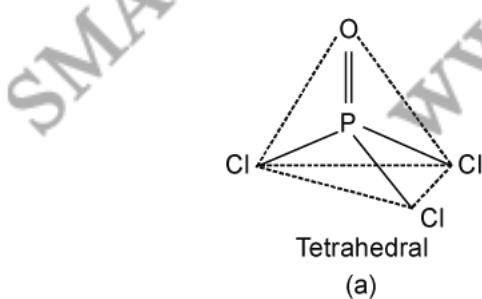
P –  — outer most electronic configuration



hybridization – sp^3d

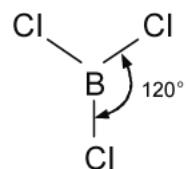
Trigonal bipyramidal

S45.

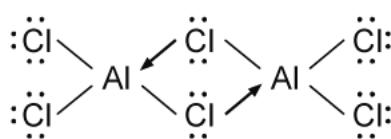


S46. In BCl_3 molecule, B atom assumes sp^2 hybrid state. The three sp^2 hybrid orbitals are used to form sigma bonds with 3 chlorine atoms to give trigonal planar structure.

Anhydrous AlCl_3 exist in the form of dimer Al_2Cl_6 . In Al_2Cl_6 each Al atom is surround by four bond pairs. As a result they assumes tetrahedral deposition.



Trigonal planar structure of BCl_3



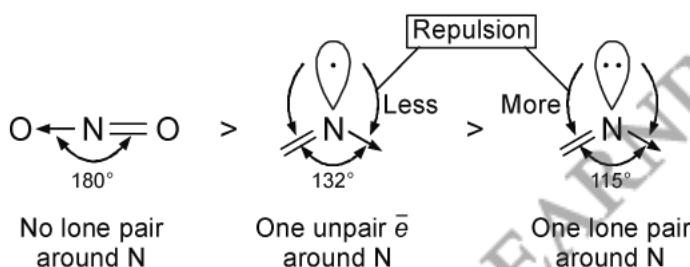
Dimer of AlCl_3

S47. (a) Lewis concept considers the formation of covalent bond by mutual sharing of electrons. VB theory considers the formation of covalent bond by overlap of half-filled atomic orbitals.

(b) Lewis concept does not provide explanation for different shapes of molecules but VB theory does explain molecular shapes.

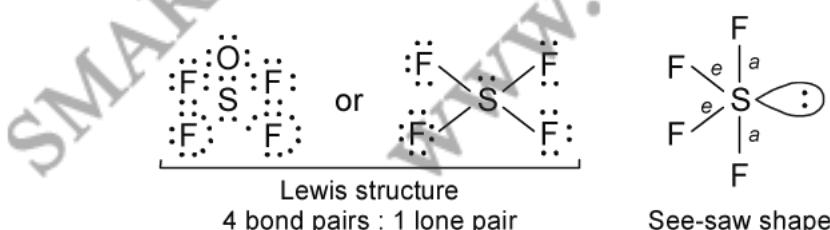
(c) Lewis concept does not explain the bond strength but VB theory is able to explain it.

S48. NO_2^+ is linear as there is no lone pair on N atom. NO_2 and NO_2^- , both have bent shape. In NO_2 , there is one odd electron on N atom but in NO_2^- there is one lone pair of electron on N atom.



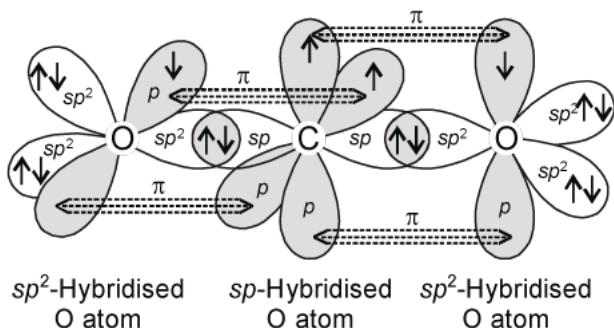
S49. Hybrid states of C atom in C_2H_6 , C_2H_4 and C_2H_2 are respectively sp^3 , sp^2 and sp . C—H bonds in C_2H_6 , C_2H_4 and C_2H_2 are respectively $sp^3\text{-s}\sigma$, $sp^2\text{-s}\sigma$ and $sp\text{-s}\sigma$ bond. Since size of hybrid orbitals varies as $sp^3 > sp^2 > sp$. Therefore, $sp^3\text{-s}\sigma > sp^2\text{-s}\sigma > sp\text{-s}\sigma$ bond.

S50. The Lewis formula of SF_4 molecule indicates the presence of one lone pair and four bond pairs. The five electron pairs around S atom adopts *trigonal bipyramidal geometry*. In this arrangement, the electron pairs at the equatorial positions experience relatively lesser repulsive interactions as compared to axial relatively lesser repulsive interactions as compared to axial electron pairs. Hence the lone pair tends to occupy equatorial position. Thus, the shape of the molecule SF_4 is just like see-saw.



S51. In CO_2 molecule, hybrid state of C atom is sp and that of each O atom is sp^2 .

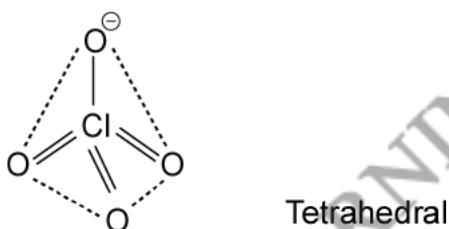
Two sp -hybrid orbitals are used to form $sp-sp^2\sigma$ bonds with two oxygen atoms. Unhybridised $2p$ -orbitals of C atom forms π -bonds lie in perpendicular planes. The molecule is **linear**. The orbital overlap in CO_2 is shown as follows.



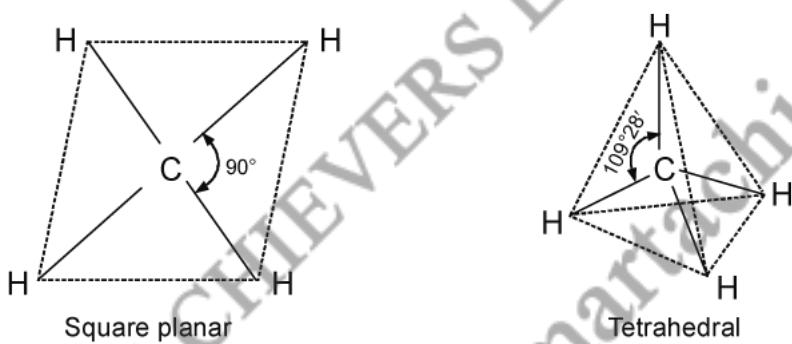
S52. NO_3^- Hybridization = sp^2



ClO_4^- Hybridization = sp^3



S53. The tetrahedral and square planar structures of CH_4 are shown:



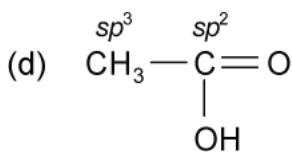
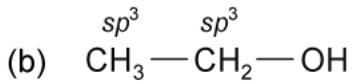
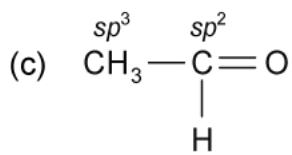
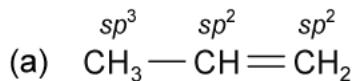
According to VSEPR theory, the electron pairs around the central atom lie as far apart as possible so as to minimise the repulsive interactions between them. For square planar arrangement for CH_4 , the four C — H bond pairs have to lie at an angle of 90° .

But in tetrahedral arrangement, the bond pairs of C — H bonds lie at an angle of $109^\circ 28'$ and therefore, have smaller repulsive interactions than those in square planar arrangement.

S54.

- (a) $\text{CH}_3 - \text{CH} = \text{CH}_2$
- (b) $\text{CH}_3 - \text{CH}_2 - \text{OH}$
- (c) $\text{CH}_3 - \overset{\text{sp}^3}{\underset{\text{H}}{\text{C}}} = \text{O}$
- (d) $\text{CH}_3 - \overset{\text{sp}^3}{\underset{\text{OH}}{\text{C}}} = \text{O}$

S55. During combination of species BF_3 and NH_3 , N atom is donor and B atom of BF_3 is acceptor. The hybrid state of B in BF_3 is sp^2 and that of in NH_3 is sp^3 . In the compound $\text{F}_3\text{H}^+ - \text{NH}_3^-$ both N and B atoms are surrounded by four bond pairs. Thus, the hybrid state of both is sp^3 . Hence, during the reaction the hybrid state of B changes from sp^2 to sp^3 but that of N it remains the same.



S56. In AlCl_3 ,

$$X = 3 + \frac{1}{2} (3 - 3 + 1) = 3$$

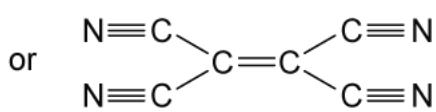
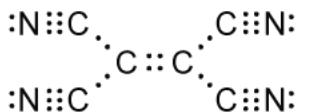
[\because Hybrid state of Al = sp^2]

In $[\text{AlCl}_4]^-$,

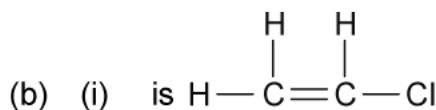
$$X = 4 + \frac{1}{2} (3 - 3 + 1) = 4$$

[\because Hybrid state of Al = sp^3]

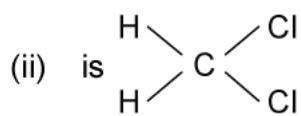
S57. (a) The Lewis structure is



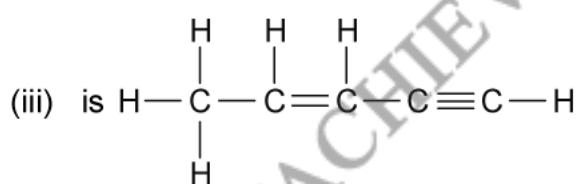
Since each triple bond has one σ and two π -bonds and each double bond has one σ and one π -bond. Thus total number of **σ -bonds** in the structure is **9**. Similarly total number of **π -bonds** is also **9**.



It contains 5 σ -bonds and 1 π -bond

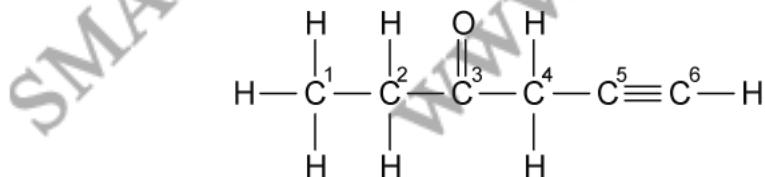


It contains only 4 σ -bonds

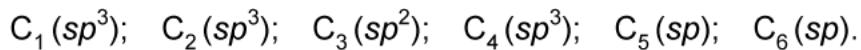


It contains 10 σ -bonds and 3 π -bonds

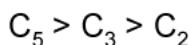
S58. The given structure is



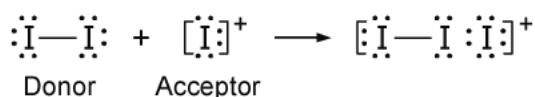
The hybrid states of various carbon atom is



(a) σ electrons in the structure = 14
 π electrons in the structure = 3
(b) C_1, C_2, C_4 atoms have sp^3 hybrid state;
 C_5, C_6 atoms have sp hybrid state
(c) Decreasing order of s -character is

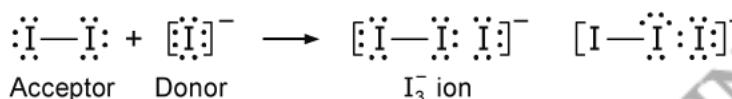


S59. The species I_3^+ is formed by the combination of I_2 and I^+ ion. They are held by coordinate bond in which I_2 molecule is donor and I^+ is acceptor.

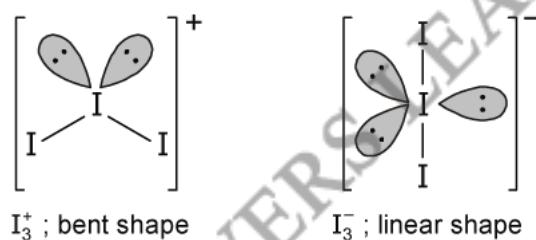


The two lone pairs and 2 bond pairs around central iodine atom adopt tetrahedral arrangement. Thus, the species has a bent shape.

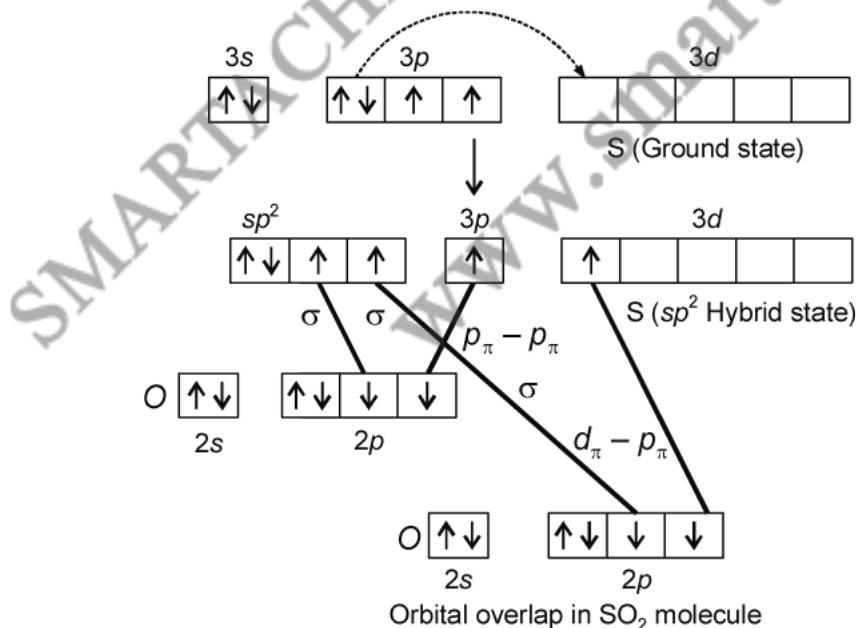
I_3^- on the other hand is formed by combination of I_2 and I^- ion in which I^- ion acts as donor and I_2 molecule act as acceptor. It accommodates electrons in empty d -orbitals.



Three lone pairs and two bond pairs around central atom adopt trigonal bipyramidal arrangement and thus, it has linear shape.

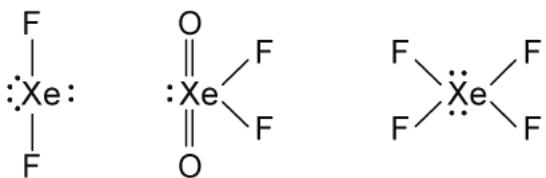


S60. Sulphur atom in SO_2 molecule assumes sp^3 hybrid state.

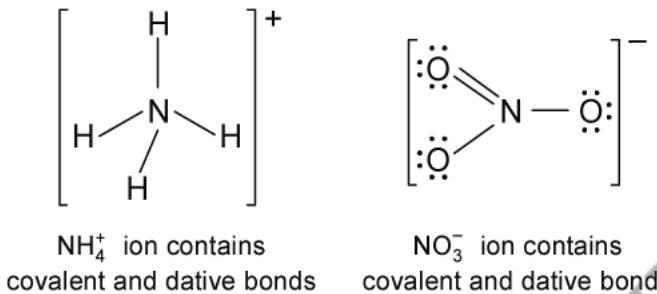


The two sp^2 -hybrid orbitals of S atom form (sp^2-p) σ bonds with two O atoms. The unhybridised $3p$ -orbital of S atom forms $p\pi-p\pi$ bond with one O atom and d -orbital of sulphur is used to form $d\pi-p\pi$ bond with second O atom.

S61. The electron dot structures of various species are as follows:

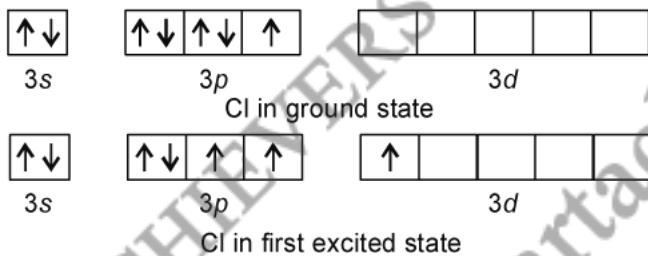


S62. NH_4NO_3 is ionic compound in which the cationic species is NH_4^+ ion and anionic species is NO_3^- ion. NH_4^+ ion is formed by the combination of NH_3 molecule and H^+ ions through dative bond. The hybrid state of N in NH_4^+ ion is sp^3 and the species is **tetrahedral**. In NO_3^- ion the N atom assumes sp^2 hybrid state and the species is **planar**.



Thus, NH_4NO_3 contains ionic bonds covalent bonds and co-ordinate bonds.

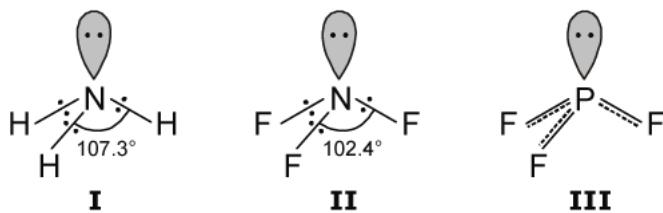
S63. Cl atom has empty d -orbitals and it acquires excited state at the time of bonding when electrons from $3p$ -orbitals are promoted to $3d$ -orbitals.



In first excited state Cl atom can exhibit a covalency of three. Hence, ClF_3 is possible. F atom cannot expand its octet due to absence of empty d -orbitals in 2nd energy shell. Hence, it cannot exhibit covalency more than 1. Therefore, FCl_3 is not possible.

S64. Both NH_3 and NF_3 have pyramidal shape with one lone pair on N atom.

However as F has higher electronegativity than H, the electron pair is attracted more towards F in NF_3 i.e., the bond pairs of electrons are away from N or in other words, distance between bond pairs is more. Hence repulsive interactions between the bond pairs in NF_3 are less than those in NH_3 . As a result, the bond angle in NF_3 is smaller (102.4°) than that in NH_3 (107.3°).



PH_3 and PF_3 are also pyramidal in shape with one lone pair on P. But PF_3 has greater bond angle than PH_3 . This is due to resonance in PF_3 leading to partial double bond character which arises due to back π -donation from F atom to empty d orbitals of P as shown in structure III. As a result, repulsions between P — F bond are larger than those between P — H bonds. Hence, the bond angle is large in PF_3 . There is no possibility back π -donation in PH_3 .

S65. (a) $\text{C}_2\text{H}_2(180^\circ) > \text{BCl}_3(120^\circ) > \text{CH}_4(109^\circ.28') > \text{NH}_3(107^\circ) > \text{H}_2\text{O}(104.5^\circ)$
 (b) $\text{NH}_4^+(109^\circ.28') > \text{NH}_3(107^\circ) > \text{NH}_2^-(\approx 104.5^\circ)$
 (c) $\text{CO}_2(180^\circ) > \text{H}_2\text{O}(104.5^\circ) > \text{H}_2\text{S}(92^\circ) > \text{H}_2\text{Se}(90.5^\circ) > \text{H}_2\text{Te}(\approx 90^\circ)$

Q1. Use Hanney Smith equation to calculate percentage ionic character in (i) H — F (ii) Si — H bond.

Q2. Dipole moment of KCl is 3.336×10^{-29} Cm. The interatomic distance between K^+ and Cl^- ions in this molecule is 2.6×10^{-10} m. Calculate the percent ionic character in KCl.

Q3. Although both CO_2 and H_2O are triatomic molecules, the shape of H_2O molecule is bent while that of CO_2 is linear. Explain this on the basis of dipole moment.

Q4. Out of H_2O and H_2S which is more polar?

Q5. In which of the following has higher dipole moment 1-butene and 1-butyne why?

Q6. CO_2 is non-polar while H_2O is polar. What conclusion do you draw about their structures from these?

Q7. Arrange the following molecules in order of increasing ionic character of their bonds: LiF , K_2O , N_2 , SO_2 , ClF_3

Q8. Arrange the following bonds in the order of increasing ionic character:
 $C - H$, $F - H$, $Br - H$, $Na - I$, $K - F$ and $Li - Cl$

Q9. Which of the two molecules OCS and CS_2 has a higher dipole moment and why?

Q10. Arrange the following bonds in order of increasing polarity:
 $P - H$, $H - O$, $C - Cl$

Q11. Arrange the following according to decreasing ionic character
 $C - H$, $F - H$, $Na - F$, $Na - I$, $K - F$.

Q12. Arrange the following according to decreasing dipole-moment
 $C - H$, $H - H$, $S - H$, $O - H$, $N - H$

Q13. Arrange the following according to decreasing dipole-moment
 CO_2 , CCl_4 , SO_2 , H_2O

Q14. Arrange the following according to decreasing dipole-moment
 H_2 , HF , HCl , HBr , HI .

Q15. If electronegativities of H and F are given to be 2.1 and 4.0 respectively, what is the % ionic character in H — F bond on the basis of Hannay Smith equation?

Q16. Why dipole-moment BF_3 , and CCl_4 are zero.

Q17. The percentage ionic character in certain bond $A - B$ is 76.81% and the bond length $A - B$ is given as 159.6 pm. What is the dipole moment of AB molecule?

Q18. Compute the partial positive charge on H-atoms in two molecules $H - Cl$ and $H - I$ respectively. Given at $\mu HCl = 1.03 D$ and $\mu(H - I) = 0.38 D$. Also $H - Cl$ and $H - I$ bond distance are 127 pm and 161 pm respectively.

Q19. (a) Calculate the dipole moment of HCl by assuming a complete transfer of an electron from H atom to Cl atom. The H — Cl bond length is 127 pm.

(b) The H — Cl bond is covalent polar and its dipole moment is 1.0 debye. From this information compute the ionic character of H — Cl bond.

Q20. Dipole-moment of BeF_2 is zero given quantitative. Justification μ .

Q21. Which out of NH_3 and NF_3 has higher dipole moment and why?

Q22. BF_3 is a nonlinear molecule and each B — F bond has a finite bond moment. But the dipole moment of BF_3 is zero.

Q23. Why dipole-moment PF_5 and SF_6 is zero.

S1. According to Haney Smith equation, % ionic character = $16(\Delta\chi) + 3.5(\Delta\chi)^2$

(i) H — F bond = $16 \times 1.9 + 3.5(1.9)^2 = 43.03\%$.

(ii) Si — H bond = $16 \times (0.3) + 3.5(1.9)^2 = 17.23\%$.

S2. Consider $K^+ Cl^-$ to be 100% ionic, then

$$\text{Charge} \times \text{distance} = 1.602 \times 10^{-19} \times 2.6 \times 10^{-10} \\ = 4.1652 \times 10^{-29} \text{ Cm}$$

$$\text{Percent ionic character} = \frac{3.336 \times 10^{-29}}{4.1652 \times 10^{-29}} = 0.8009 \text{ or } 80.09\%.$$

S3. The bond moments of two C = O bonds in CO_2 cancel each other indicating the linear structure for CO_2 . The resultant μ for H_2O molecule is $\neq 0$. This indicates bent structure for H_2O molecule because bond moments of two O — H bonds do not cancel out.



S4. H_2O is more polar than H_2S because of oxygen is more electronegative than sulphur and H_2O bond angle is higher than H_2S



S5. $CH_3 — CH_2 — CH = CH_2$ is dipole moment is lower than $CH_3 — CH_2 — C \equiv CH$, because of in butene $= C$ is sp^2 hybridized which is less electronegative while $\equiv C$ is sp hybridized which is more electronegative.

S6. CO_2 is a linear molecule, therefore the resultant dipole moment of two C = O bonds get cancelled giving zero dipole moment. On the other hand, water molecule is not linear. It has angular shape and bond moments of two O — H bonds resultant add and molecule become polar.



S7. $N_2 < SO_2 < CIF_3 < K_2O < LiF$

S8. $C — H < Br — H < F — H < Li — Cl < Na — I < K — F$

S9. Both OCS and CS_2 are linear molecules having, Lewis structure as



In CS_2 , the dipole moment is 0 because the bond moments cancel out. But in OCS, bond moment of C = O is not equal to that of C = S. Hence, it has a net dipole moment. Thus, dipole moment of OCS is higher.

S10. Order of polarity is : P — H < C — Cl < H — O

S11. K — F > Na — F > Na — I > H — F > C — H.

S12. O — H > N — H > S — H > C — H > H — H.

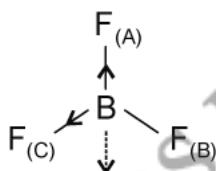
S13. $\text{H}_2\text{O} > \text{SO}_2 > \text{CO}_2 > \text{CCl}_4$.

S14. HF > HCl > HBr > HI > H_2

S15. $\chi_{\text{H}} = 2.1$; $\chi_{\text{F}} = 4.0$

$$\begin{aligned} \text{Now,} \quad \% \text{ of ionic character} &= 16(4 - 2.1) + 3.5(4 - 2.1)^2 \\ &= 16 \times 1.9 + 3.5(19)^2 \\ &= 30.4 + 12.635 = \mathbf{43.035\%}. \end{aligned}$$

S16. BF_3 is triangular planar and resultant dipole-moment of two B — F bond is equal and opposit in direction



(Resultant of B — C is equal and opposit to A)



CCl_4 is square planar. So, dipole moment of each C — Cl bond is canceled out. So net dipole moment is zero.

S17. If the molecule were 100% ionic, then

$$\begin{aligned} \mu_{\text{ionic}} &= q \times d \\ &= 1.602 \times 10^{-19} (\text{C}) \times 159.6 \times 10^{-12} (\text{m}) \\ &= 2.557 \times 10^{-29} \text{ Cm} \end{aligned}$$

$$\text{Now,} \quad \% \text{ ionic character} = \frac{\mu_{\text{obs}}}{\mu_{\text{ionic}}} \times 100$$

$$\text{or} \quad \mu_{\text{obs}} = \frac{\% \text{ ionic character} \times \mu_{\text{ionic}}}{100}$$

$$= \frac{76.81 \times 2.557 \times 10^{-29}}{100} \\ = 1.96 \times 10^{-29} \text{ Cm}$$

$$= \frac{1.96 \times 10^{-29}}{3.335 \times 10^{-30}} D = 5.87 D.$$

S18. Let δ be the magnitude of positive charge on H atom

For HCl

$$\delta = \frac{\mu_{\text{HCl}}}{d} = \frac{1.03 \times 10^{-18}}{127 \times 10^{-10}} = 8.11 \times 10^{-11} \text{ esu}$$

For H — I

$$\delta = \frac{\mu_{\text{HI}}}{d} = \frac{0.38 \times 10^{-18}}{61 \times 10^{-10}} = 2.4 \times 10^{-11} \text{ esu}$$

Now,

$$\frac{\delta_{\text{HCl}}}{\delta_{\text{HI}}} = \frac{8.11 \times 10^{-11}}{2.4 \times 10^{-11}} = 3.3 : 1.$$

S19. (a) In case of complete transfer of electron the charge on each end of the dipole should be unity.

That is, charge is equal to 1.602×10^{-19} C. Thus

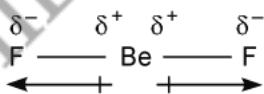
$$\begin{aligned} \mu(\text{Theoretical}) &= \text{Theoretical charge} \times \text{Bond length} \\ &= 1.602 \times 10^{-19} \text{ C} \times 127 \times 10^{-12} \text{ m} \\ &= 2.03 \times 10^{-29} \text{ C m} = 6.08 \text{ D} \end{aligned}$$

(b) $\mu(\text{actual}) = 1.0 \text{ D} = 3.335 \times 10^{-30} \text{ C m}$

$$\text{Ionic character} = \frac{\mu(\text{actual})}{\mu(\text{theoretical})} = \frac{1.0 \text{ D}}{6.08 \text{ D}} = 0.164 = 16.4\%$$

Percentage ionic character of H — Cl bond = 16.4%

S20. BeF_2 has two Be-Fe bond in equal and opposite direction bond moment.



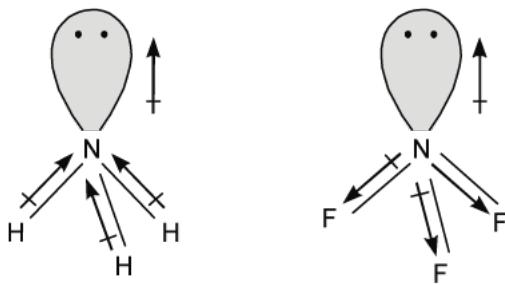
The resultant dipole moment (μ) of a molecule is

$$\mu^2 = \mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos 180$$

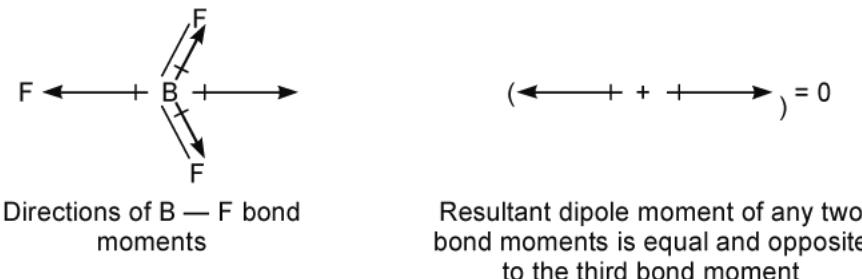
$$\therefore \theta = 180 \text{ and } \cos 180 = -1$$

$$\therefore \mu^2 = \mu_1^2 + \mu_2^2 + 2\mu_1^2 (-1) = 2\mu_1^2 - 2\mu_1^2 = 0$$

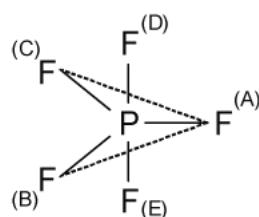
S21. NH_3 has higher dipole moment than NF_3 . Both NH_3 and NF_3 molecules have pyramidal shape with one lone pair on nitrogen atom. In case of NH_3 , the orbital dipole due to lone pair is in the same direction as the resultant dipole moment of the three N — H bonds. Therefore, it adds on the resultant dipole moment of the N — H bonds. On the other hand, in case of NF_3 , the orbital dipole is in the opposite direction to the resultant dipole moment of the three N — F bond moments as shown in figure. Consequently, the dipole moment of NF_3 is low.



S22. The electronegativity of F is more than that of B, so a polarity is developed in each B — F bond. But the orientation of each B — F bond is such that the resultant dipole moment of any two B — F bonds vectorially cancels the dipole moment of the third B — F bond. Thus, the dipole moment of BF_3 is zero.



S23. PF_5 is triangular pyramidal shape. Having three P — F arrange in 120° angle.



$\text{F}_C — \text{F}_B$ resultant is equal and opposit to F_A , and $\text{F}_D — \text{F}_E$ are opposit. i.e., resultant dipole moment is zero.

SF_6 is octahedral in shape, all S — F bond are 90° and neutriliye each other

