Chapter: 06

Chemical Bonding

Objective

1. Dipole Moment of H_2O is:
Ⓐ 0.12 D Ⓑ 0.95 D Ⓒ 1.61 D [✓ Ⓓ 1.85 D]

2. The pH of 10^{-3} mol dm $^{-3}$ of an aqueous solution of H_2SO_4 is:
Ⓐ 1.5 Ⓑ 2.0 [✓ Ⓒ 2.7] Ⓓ 3.0

3. In nitrogen molecule (N_2), each nitrogen atom contributes in sharing for formation of bond:
Ⓐ four electrons [✓ Ⓑ three electrons] Ⓒ two electrons Ⓓ one electron

4. Which one has highest value of ionization energy:
[✓ Ⓐ F] Ⓑ Be Ⓒ C Ⓓ O_2

5. Geometry of SO_2 molecule is:
Ⓐ Trigonal pyramidal Ⓑ Linear Ⓒ Tetrahedral [✓ Ⓓ Angular]

6. 1 A° = ……. m.
Ⓐ 157 torr Ⓑ 158 torr [✓ Ⓒ 10^{-10}] Ⓓ 156 torr

7. Which of the following molecules has zero dipole moment?
Ⓐ $CHCl_3$ [✓ Ⓑ CS_2] Ⓒ NH_3 Ⓓ H_2O

8. Optimum temperature for synthesis of ammonia by Haber Process is:
[✓ Ⓐ 400 °C] Ⓑ 410 °C Ⓒ 390 °C Ⓓ 370 °C

9. The type of hybridization in $BeCl_2$ is:
Ⓐ dSp^2 Ⓑ Sp^3 Ⓒ Sp^2 [✓ Ⓓ sp]

10. Minimum amount of energy required to remove an electron from its gaseous atom is called:
Ⓐ Reduction Ⓑ Oxidation [✓ Ⓒ Ionization energy] Ⓓ Electron-Affinity

11. The amount of energy released by absorbing electron in the valence shell is:
Ⓐ Atomization energy [✓ Ⓑ Electron affinity]
Ⓒ Electronegativity Ⓓ Ionization energy

12. Most stable electronic configuration is of a/an:
Ⓐ Alkali metal [✓ Ⓑ Noble gas] Ⓒ Halogen Ⓓ Electronegative element

13. The covalent radius of Cl-atom is:
[✓ Ⓐ 99.4 pm] Ⓑ 66.4 pm Ⓒ 70 pm Ⓓ 80 pm

14. The bond order of He_2 is:
Ⓐ One Ⓑ Two Ⓒ Three [✓ Ⓓ Zero]

15. In sp hybrid orbital percentage of S-character is:
Ⓐ 25% [✓ Ⓑ 50%] Ⓒ 75% Ⓓ 100%

16. In sp^3 hybrid orbital percentage of S-character is:

Ⓐ 50% [✓ Ⓑ 25%] Ⓒ 75% Ⓓ 100%

17. Bond order of N_2 molecule is:

Ⓐ 04 [✓ Ⓑ 03] Ⓒ 02 Ⓓ 01

18. The electron affinity of chlorine is:

Ⓐ $+396$ KJ mol^{-1} Ⓑ -449 KJ mol^{-1} Ⓒ -249 KJ mol^{-1} [✓ Ⓓ -349 KJ mol^{-1}]

19. Noble gases are highly stable and least reactive because:

[✓ Ⓐ Their valence shell are complete] Ⓑ They are very safe

Ⓒ They are present in zero group Ⓓ They are gasses

20. Octet rule is not obeyed during its formation:

Ⓐ CCl_4 [✓ Ⓑ PCl_5] Ⓒ CF_4 Ⓓ NF_3

21. Which of the following Molecule Obey Octet Rule:

[✓ Ⓐ NF_3] Ⓑ SF_6 Ⓒ BCl_3 Ⓓ BF_3

22. Molecule in which the distance between two carbon atoms is the largest is:

Ⓐ C_2H_2 Ⓑ C_2H_4 Ⓒ C_6H_6 [✓ Ⓓ C_2H_6]

23. The radius of Na^{+1} ion is:

Ⓐ 93 pm [✓ Ⓑ 95 pm] Ⓒ 94 pm Ⓓ 92 pm

24. Total number of bonds in C_2H_4 molecule are:

Ⓐ Eight Ⓑ Five [✓ Ⓒ Six] Ⓓ Four

25. The number of bond in oxygen molecule is:

Ⓐ Two sigma only [✓ Ⓑ One σ and one π] Ⓒ One σ and two π Ⓓ Three sigma only

26. The number of bonds in nitrogen molecule are:

Ⓐ Two sigma only [✓ Ⓑ One σ and two π] Ⓒ One σ and one π Ⓓ Three sigma only

27. In methanol, bond between carbon and oxygen is:

[✓ Ⓐ Polar] Ⓑ Co-ordinate Ⓒ Ionic Ⓓ Nonpolar

28. CsF has ionic character:

Ⓐ 100% Ⓑ 80% Ⓒ 70% [✓ Ⓓ 92%]

29. The most electronegative element is:

Ⓐ Hydrogen [✓ Ⓑ Fluorine] Ⓒ Oxygen Ⓓ Nitrogen

30. The shielding effect is responsible for:

Ⓐ The increase in nuclear attractive influence over the valence electrons.

[✓ Ⓑ The decrease in nuclear attractive influence over the valence electrons.]

Ⓒ The increase in attraction between nucleus and inner electrons.

Ⓓ The decrease repulsion between nucleus and inner electrons.

31. The value of third ionization energy of Mg is:

Ⓐ 7850 kJ mol^{-1} Ⓑ 1890 kJ mol^{-1} [✓ Ⓒ 7730 kJ mol^{-1}] Ⓓ 1450 kJ mol^{-1}

32. First ionization Energy of Mg atom is:

[✓ Ⓐ $+738$ KJ mol^{-1}] Ⓑ -500 KJ mol^{-1} Ⓒ $+1450$ KJ mol^{-1} Ⓓ -349 KJ mol^{-1}

33. Ionic and co-ordinate covalent bonds are present in:
Ⓐ C_2H_5 [✓ Ⓑ NH_4CI] Ⓒ H_2O Ⓓ SO_2

34. Which of the following has coordinate covalent bond?
Ⓐ $AlCl_3$ Ⓑ HCl [✓ Ⓒ NH_4CI] Ⓓ NaCl

35. The molecules which cannot form co-ordinate covalent bond with H^+ ion is:
Ⓐ PH_3 [✓ Ⓑ CH_4] Ⓒ NH_3 Ⓓ H_2O

36. Who developed the VSEPR theory:
Ⓐ Bolzmann Ⓑ Bernoulli Ⓒ Maxell [✓ Ⓓ Nylholm and Gillespie]

37. Which of the following has linear structure?
Ⓐ H_2O [✓ Ⓑ CO_2] Ⓒ CH_4 Ⓓ NH_3

38. Which of the following has bond angle of 120°:
[✓ Ⓐ BF_3] Ⓑ CH_4 Ⓒ $BeCl_2$ Ⓓ NH_3

39. The geometry of ethane is:
Ⓐ Linear [✓ Ⓑ Tetrahedral] Ⓒ V – shaped Ⓓ Trigonal planar

40. Which of the hydrogen halides has the highest percentage of ionic character?
[✓ Ⓐ HF] Ⓑ HBr Ⓒ HI Ⓓ HCI

41. The bond angle in NH_3 molecule is:
Ⓐ 104.5° Ⓑ 108.5° Ⓒ 109.5° [✓ Ⓓ 107.5°]

42. The molecular shape of SO_3 is:
Ⓐ Tetrahedral Ⓑ Linear [✓ Ⓒ Triangular planar] Ⓓ Pyramidal

43. Shape of $SnCl_2$ is:
Ⓐ Tetrahedral [✓ Ⓑ bent or angular] Ⓒ trigonal (planar) Ⓓ Linear

44. The hybridization in ammonia molecule is:
[✓ Ⓐ sp^3] Ⓑ sp^2 Ⓒ sp Ⓓ dsp

45. Carbon atom in CH_4 is hybridized:
Ⓐ dsp^2 Ⓑ Sp Ⓒ sp^2 [✓ Ⓓ sp^3]

46. Molecular orbital theory was proposed by:
Ⓐ Werner [✓ Ⓑ Millikan] Ⓒ Mosley Ⓓ Kossel

47. An ionic compound $A^+ B^-$ is most likely to be formed when:
Ⓐ both the ionization energy of A and electron affinity of B are high
[✓ Ⓑ the ionization energy of A is low and electron affinity of B is high]
Ⓒ both the ionization energy of A and electron affinity of B are low
Ⓓ the ionization energy of A is high and electron affinity of B is low

48. Which of the following has the highest bond order?
Ⓐ O_2^{-2} [✓ Ⓑ O_2^{-1}] Ⓒ O_2^{+2} Ⓓ O_2^{+1}

49. Bond order of O_2^{-2} is:

[✓ Ⓐ One] Ⓑ Two Ⓒ Three Ⓓ Zero

50. Which of the following species has unpaired electrons in ant bonding molecular orbitals?

Ⓐ F_2 Ⓑ Br_2 Ⓒ O_2^{+2} [✓ Ⓓ N_2^{-2}]

51. Which of the following molecule has zero dipole moment:

Ⓐ H_2O [✓ Ⓑ BF_3] Ⓒ NH_3 Ⓓ $CHCl_3$

52. The Sl unit of dipole moment is:

[✓ Ⓐ mC] Ⓑ Nm^{-2} Ⓒ Debye Ⓓ Joule

53. The dipole moment of SO_3 is:

Ⓐ 0.95 D Ⓑ 0.78 D Ⓒ 0.12 D [✓ Ⓓ 0.0 D]

54. The dipole moment of SO_3 is:

Ⓐ 0.95 D Ⓑ 1.49D [✓ Ⓒ 1.61D] Ⓓ 1.85 D

55. ………… molecules have zero dipole moment.

[✓ Ⓐ CH_4] Ⓑ SO_2 Ⓒ CO Ⓓ H_2S

56. Dipole moment of CO_2 is:

Ⓐ 3.1D Ⓑ 1.85D [✓ Ⓒ zero] Ⓓ 1.25D

57. Which one of the given compounds possesses ionic bonding?

Ⓐ CH_3CI [✓ Ⓑ CaO] Ⓒ CH_2H_6 Ⓓ CH_4

58. How many types of bonds in NH_4Cl are:

Ⓐ Four types [✓ Ⓑ Three types] Ⓒ Two types Ⓓ One type

59. Ionization energy for $Mg \rightarrow Mg^{+} + 1e^{-}$ has $\Delta H = ?$

[✓ Ⓐ 738 KJ mol^{-1}] Ⓑ 138 KJ mol^{-1} Ⓒ 238 KJ mol^{-1} Ⓓ 438 KJ mol^{-1}

60. Which one of the following hydrocarbon has shortest C-C bond length?

Ⓐ Benzene Ⓑ Ethane Ⓒ Ethene [✓ Ⓓ Ethyne]

61. Ethyne molecule have:

Ⓐ One π and two σ bonds [✓ Ⓑ One σ bonds between carbon atom]
Ⓒ Three σ bonds between carbon atom Ⓓ Three π bonds between carbon atom

62. According to VSEPR theory, the shape of PH_3 molecule is:

Ⓐ Trigonal planar Ⓑ Linear [✓ Ⓒ Trigonal Pyramidal] Ⓓ Tetragonal

63. An orbital which is spherical and symmetrical is:

[✓ Ⓐ s − orbital] Ⓑ f − orbital Ⓒ d −orbital Ⓓ p − orbital

64. Which molecule has sp^2 hybridization?

Ⓐ C_2H_6 [✓ Ⓑ C_2H_4] Ⓒ CH_4 Ⓓ C_2H_2

65. The Carbon atom in C_2H_4 is:

[✓ Ⓐ sp^2 − hybridized] Ⓑ Sp^3 − hybridized Ⓒ dsp^2 − hybridized Ⓓ sp − hybridized

Fill in the blanks

Q1: The tendency (UB) of atoms to attain maximum of .... electrons in the valence shall is called completion of octet.

Q2: The geometrical shape of $SiCl_4$ and PCl_3 can be explained on the basis of hybridizations.

Q3: For N_2 molecule the energy of (2px) orbital is .... than π (2py) orbital.

Q4: The paramagnetic property of O_2 is well explained on the basis of MO theory in terms of the presence of .... electrons in two MO orbitals.

Q5: The bond order of N_2 .... while that of Ne_2 is ....

Q6: The values of dipole moment of CS_2 is while for SO_2 is ....

Q7: The VSEPR theory stands for ....

Answers

1.	8	2.	Sp^3	3.	greater	4.	unpaired
5.	3, zero	6.	Zero, non-zero	7.	valence shell electron pair repulsion		

Chapter : 06 Chemical Bonding

Subjective

Q1: ***Why Noble gases don't form chemical bonds?***

Ans: *Noble gases have two or eight electrons in their valence shell so they have completely filled their valence shell. They do not have vacant space to accommodate extra electrons. This is the reason that noble gasses do not lose, gain or share electrons and do not form chemical bond.*

Q2: ***Define octet rule. Give two examples of compounds which do not obey this rule.***

Ans: *Octet rule:*

The tendency of atoms to attain a maximum of eight electrons in the valence shell is known as Octet rule. Many compounds do not obey this rule.

For example:

PCl_5, BF_3, SF_6 *etc.*

Q3: ***What is the cause for chemical change?***

Ans: *It is the natural tendency of each element in periodic table to attain the electronic configuration of nearest noble gas by:*

- *Gaining electrons*
- *Sharing of electrons*
- *Losing electrons*

Hence chemical bonds are formed.

Q4: ***Why the atoms of the elements other than noble gases combine with one another?***

Ans: *Atoms can combine with each other to attain the 2 or 8 electrons in their valence shell. That's why they combine with each other. While noble gases have completed their octet.*

Q5: ***Define bond energy with two parameters which determine its strength?***

Ans: *Bond energy:*

The amount of energy required to break 6.02×10^{23} *bonds of particular type is called bond energy.*

Its unit is $\mathrm{KJ\ mole^{-1}}$. *It depends upon following factors.*

- *Size of atoms.*
- *Bond length.*
- *Electronegativity difference between bonded atoms.*

Q6: ***Compare bond strength of polar and non−polar molecules?***

Ans: *Polar covalent bond is stronger than non-polar covalent bond. Bond energy generally increases due to increases in electronegativity difference between bonded atoms and it is because of increase in percentage ionic character.*

Q7: ***Why the covalent compounds show isomerism, but ionic compounds do not?***

Ans: *The compounds which have same molecular formula but different structures show the phenomenon of isomerism. Ionic compound have formula unit not have molecular formula that's why ionic compounds do not show isomerism.*

Q8: ***Write short note on atomic radii and ionic radii. Give example.***

Ans: *Atomic radius:*

The one-half distance between the nuclei of two alike bonded atoms is called atomic radius.

Example:

Atomic radius of Na *is* 186 pm.

Ionic radius:

The radius of an ion while considering it spherical is called ionic radius.

Example:

Ionic radius of Na *is* 95.

Q9: ***Why atomic radius is greater than cationic radius?***

Ans: *When an atom lose one or more electrons to form cation. In positive ions number of electrons are reduced but positive charge on nucleus remain same. Therefore nucleus attracts electrons inward resulting in decrease of ionic radius.*

Hence positive ions always have smaller size than neutral atom.

Q10 ***Why cationic radii are smaller than anionic radii?***

Ans: *The ionic radius of a cation is smaller than the atomic radius of the element from which it is derived. The ionic radius of an anion is greater than the atomic radius of the corresponding atom. This is due to the reason, that with the successive loss of electrons the nuclear charge attracts the remaining electrons with a greater force the increase in the size of anion is due to the increase in electron-electron repulsion because of the increase in the valence shell electrons.*

Q11: ***Why cationic radius is smaller than parent atom?***

Ans: *The ionic radius of a cation is smaller than the atomic radius of the element from which it is derived. This is due to the reason that with the successive loss of electrons the nuclear charge attracts the remaining electrons with a greater force.*

Q12: ***Define electronegativity and electron affinity of an atom?***

Ans: *Electronegativity:*

The tendency of an atom to attract shared pair of electron toward itself is called electronegativity.

Electron Affinity:

The amount of energy released when an electron is added to an empty or partially filled orbital of an atom in its valence shell to form an anion is called electron affinity.

For example:

$$Cl + e^- \longrightarrow Cl^- \qquad \Delta H = -349\ KJ\ mol^{-1}$$

Q13: ***Name the factors influencing the electron affinity.***

Ans: *Factors influencing the electron Affinity:*

- *Nuclear charge*
- *Atomic radius*
- *Shielding effect of inner shell electrons.*

Q14: ***How electron affinity changes in a group?***

Ans: *In groups the atomic radii increase with the increase in the proton number due to Successive increase electronic shells. This also exerts a shielding effect on the force of attraction between the nucleus and the valence electrons. Thus, the electron affinities usually decrease from top to bottom.*

Q15: ***Fluoride has electron affinity less than the chloride. Give reason.***

Ans: *In fluorine addition of electron takes place in* 2p *orbital whereas addition of electron takes place in* 3p *orbital in chlorine.* 2p *orbital of fluorine has smaller size and thick electronic cloud that repels the incoming electron.*

But in chorine 3p *orbital has sufficient distance for incoming electron and causes to nuclear attraction decrease. Therefore, Fluoride has electron affinity less than the chloride.*

Q16: ***Why anionic radius is larger than atomic radius?***

Ans: *Negative ion is always larger in size than the neutral atom from which it is formed. It is because electron repulsions increase due to increases in number of electrons in valence shell and hold of nucleus to greater valence electrons decreases.*

For example:

Cl *Atomic radius* 99 pm

Cl^- *Atomic radius* 181 pm

Q17: ***Why the atomic radii of the atoms cannot be determined precisely?***

Ans: *Atomic radii cannot be determined precisely because*

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

Q18: ***Why the atomic radii increase down the group?***

Ans: *The increase in atomic radii in a group is due the increase in the number of shells and the screening effect.*

Q19: ***Why the atomic radius decreases along period and increases along group?***

Ans: *The atomic radii decrease from left to the right in a period and increases from top to bottom in a group of the periodic table. The decreasing trend in a period is due to the increase in the nuclear charge. As the nuclear charge increases the pull on the electrons is increased and size of an atom decreases.*

Moreover, the shielding effect remains the same from left to right in a period. The increase in atomic radii in a group is due to increase in the number of shells and the screening effect.

Q20: ***The size of chlorine Cl atom is smaller than Cl^- ion. Justify it?***

Ans: *The ionic radius of an anion is greater than the atomic radius of the corresponding atom. The increase in the size of anion is due to the increase in electron—electron repulsion because of the increase in the valence shell electrons. Hence the size of Cl^- ion increases from* 99 pm *to* 181 pm.

Q21: ***Define electronegativity. Give its trend in periodic table.***

Ans: *Electronegativity:*

The tendency of an atom to attract the shared pair of electron toward itself in a molecule is called electronegativity.

Trend in periodic table:

Electronegativity values increase from left to right in a periodic table due to decrease in atomic size. Electronegativity value decrease down the group.

Q22: ***The size of anion is greater than its parent atom. Justify.***

Ans: *The increase in the size of anion is due to the increase in electron-electron repulsion because of the increase in the valence shell electrons.*

Q23: ***Define ionization energy and electron affinity with one example in each case.***

Ans: *Ionization energy:*

The ionization energy of an. element is the minimum energy required to remove an electron from its gaseous atom to form an ion. The process is called ionization.

Example:

$$\mathrm{Mg} \longrightarrow \mathrm{Mg}^+ + 1e^- \qquad \Delta H_{i1} = +738\ \mathrm{KJ\ mol^{-1}}$$

Electron Affinity:

The electron affinity of an atom is the energy released when an electron adds to an empty or partially filled orbital of an isolated gaseous atom in its valence energy level to form an anion having a unit negative charge.

For example:

$$\mathrm{Cl_{(g)}} + e^- \longrightarrow \mathrm{Cl^-_{(g)}} \qquad \Delta H = -349\ \mathrm{KJ\ mol^{-1}}$$

Q24: ***1st ionization energy of Mg is lower than the 2nd ionization energy. Explain.***

Ans:

$$\mathrm{Mg} \longrightarrow \mathrm{Mg}^+ + 1e^- \qquad \Delta H_{i1} = +738\ \mathrm{KJ\ mol^{-1}}$$

It is because due to removal of electron, hold of the nucleus on the remaining electrons increases. Hence it is difficult to remove second electron and thus ionization energy value increase.

$$Mg^{+} \longrightarrow Mg^{+2} + e^{-} \qquad \Delta H_{i2} = +1451\ KJ\ mol^{-1}$$

Q25: ***How ionization energy does vary in periodic table?***

Ans: *In groups:*

Ionization energy decreases down the group although nuclear charge increases. It is due to increase in number of shells down the group.

In periods:

Ionization energy increase along the periods from left to right. It is due to number of shell remain same and nuclear charge increases so ionization energy also increases.

Q26: ***Ionization energy decreases down the group although atomic number is increased. Explain it.***

Ans: *It is due to increase in number of shells down the group. Due to addition of shells electrons become away from nucleus and removal of electrons become easy so ionization energy decrease down the group.*

Q27: ***Why the ionic radius of Cl^- ion increases from 99 pm to 181 pm?***

Ans: *The radius of negative ion is always larger than parent atom. It is because electron−electron repulsions. Hence the size of Cl^- ion increases from 99 pm to 181 pm.*

Q28: ***Electronegativity increase from left to right in the periodic table. Give reason.***

Ans: *Electronegativity values increases from left to right in the periodic table due to decreased in atomic size. As atomic size decreases attraction for shared electron pair increases. Hence electronegativity value increases.*

Q29: ***How electronegativity values of elements help ds in understanding the nature of chemical bond?***

Ans: *Difference in electronegativity values is an index to the polar nature of bond in a molecule. If electronegativity difference is zero, the bond is non-polar. All the bonds between different atoms are polar. If difference of electronegativity value is greater than 1.7 the bond is ionic.*

Q30: ***Ionization Energy is an Index to the metallic character. Justify it.***

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

Q31: ***Define shielding effect and how it varies along period.***

Ans: *The decrease in force of attraction between nucleus and valence electron due to presence of inner shell electrons or intervening electrons is called shielding effect.*

Shielding effect remains almost same in a period because no new shells increase between and valence shell electron.

Q32: ***How does Shielding effect affect ionization energy?***

Ans: *The force of attraction between the nucleus and the outer electron decreases due to shielding effect of the intervening electrons, so the electron can be removed more easily or with less energy from the valence shell. The value of ionization energy decreases.*

Q33: ***The difference in E.N of bonded atoms is index of polar nature of covalent bond.***

Ans: *It means when E.N difference is zero, the bond between two atoms is non−polar, while bond formed between dissimilar atoms is mostly polar. A difference of 1.7 units shows equal contributions of ionic and covalent bonds.*

Q34: ***Define Covalent bond. Give two examples.***

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

For example:

H_2 *or* Cl_2 *molecules.*

Q35: ***Define ionic and covalent radii.***

Ans: *Ionic Radii:*

The ionic radius of an ion is the radius of the ion while considering it to be spherical in shape.

For example:

Ionic radius of Na^+ *ion is* 95 pm.

Covalent Radii:

The covalent radius of an element is defined as half of the single length between two similar atoms covalently bonded in a molecule.

For example:

Covalent radius of hydrogen is 37.7 pm.

Q36: ***Differentiate between polar and non-polar covalent bond.***

Ans: *The difference between polar and non−polar covanlent bond is:*

Polar Covalent Bonds	***Non-Polar Covalent Bonds***
When two different atoms are joined by a covalent bond, the electron pair is not equally shared between the bonded atoms. The bonding pair of electrons will be displaced towards the more electronegative atom. This would make one end of the molecule partially positive and the other partially negative as shown by the following example. *Hydrogen fluoride* $H^{\delta+} - F^{-\delta}$	*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 electro negativities. Due to an even distribution of charge, the bonded atoms remain electrically neutral.* *Hydrogen* H:H or H−H

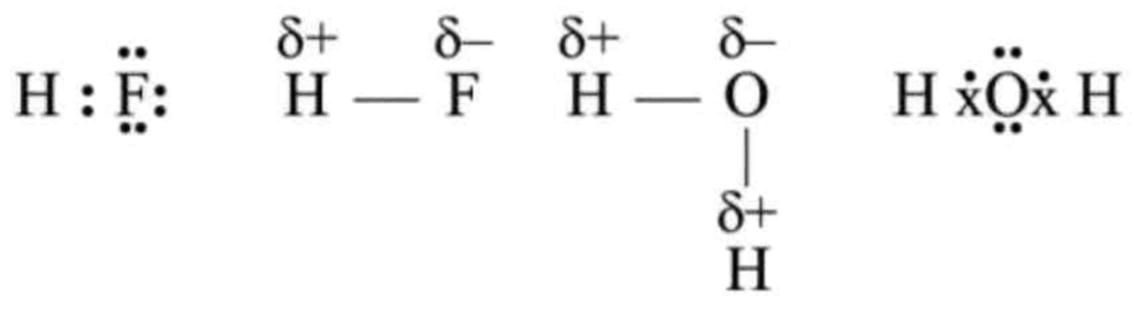
Q37: ***Define polar covalent bond with two examples.***

Ans: *Polar covalent bond:*

The covalent bond in which bonding electrons are not share equally by atoms is called polar covalent bond.

For example:

Water and Hydrogen Fluoride



Hydrogen fluoride Water

Q38: ***No bond in chemistry is 100% ionic.***

Ans: *Highest ionic character found in* CsF *is* 92%. *No metal is more electropositive than* Cs *and no non−metal is more electronegative than Fluorine, so there is no chance of* 100% *ionic character.*

Q39: ***How coordinate covalent bonds differ from covalent bond?***

Ans: *Covalent bond:*

The bond formed by mutual sharing of electrons between two atoms called covalent bond. While in coordinate covalent bond shared pair of electron donated by only one of the bonded atoms. But in some cases after bond formation distinction is present between a coordinate covalent bond and covalent bond.

For example:

H_3O^+

Q40: ***Why polar bond is stronger than non−polar bond?***

Ans: *Polar bond has* +ve *and* −ve *ends which provide extra force of attraction. Thus polar bond is stronger than non−polar bond.*

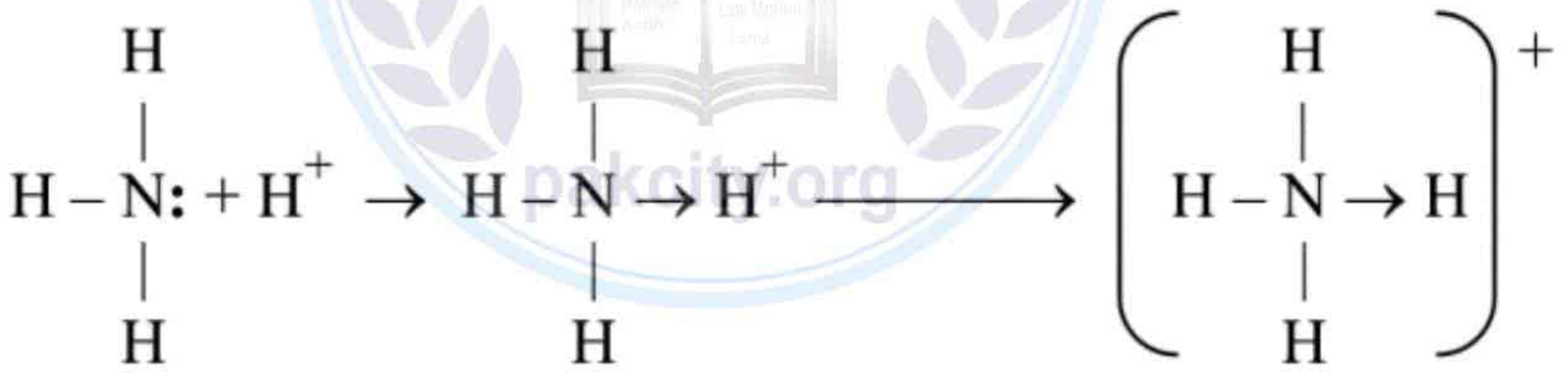
Q41: ***The distinction between a coordinate covalent bond and a covalent bond vanishes after bond formation in*** $\mathbf{NH_4^+, H_3O^+}$ ***and*** $\mathbf{CH_3NH_3^+}$.

Ans: H_3O^+:

Oxygen of water gives its lone pair to H^+ *to form coordinate covalent bond. The new bond is formed with same specie* H^+ *as two of them are already bonded. Hence all thre bonds have equal status. Every bond has* 33% *coordinate covalent character and* 66% *covalent character.*

NH_4^+:

The coordinate covalent bond is formed with same category H^+ *as already discussed. Every bond has* 25% *coordinate covalent bond character and* 75% *covalent bone character.*

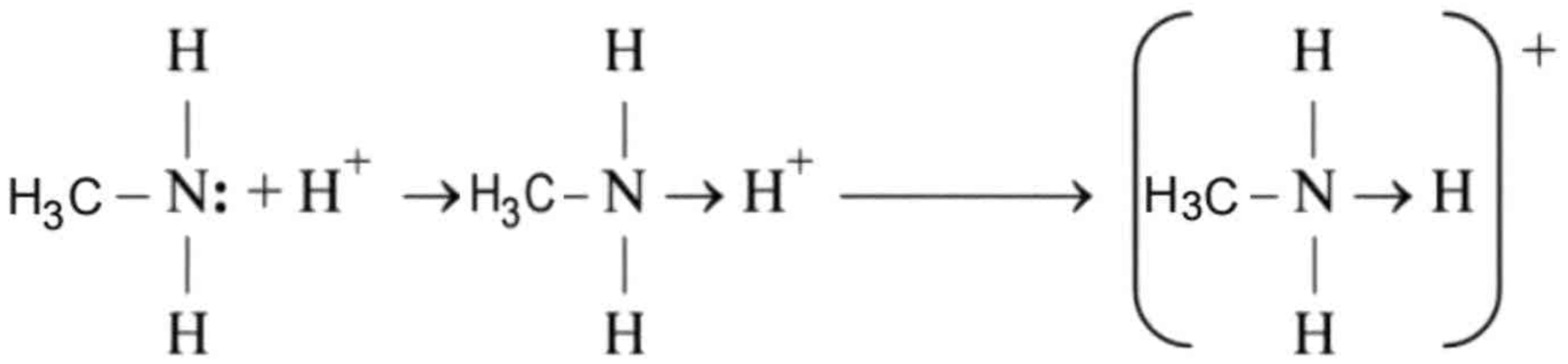


$$\mathrm{H{-}\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{:}} + H^+ \rightarrow \mathrm{H{-}\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{\rightarrow}H^+} \longrightarrow \left[\mathrm{H{-}\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{\rightarrow}H}\right]^+$$

Donor *acceptor* *ammonium ion*

$CH_3NH_3^+$:

All bonds behave alike due to same type of atoms.



$$\mathrm{H_3C{-}\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{:}} + H^+ \rightarrow \mathrm{H_3C{-}\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{\rightarrow}H^+} \longrightarrow \left[\mathrm{H_3C{-}\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{\rightarrow}H}\right]^+$$

Q42: ***Why*** $\mathbf{CO_2}$ ***and*** $\mathbf{CS_2}$ ***have linear structures?***

Ans: CO_2 *molecule is linear because all the four valence electrons of carbon atoms form double bond with oxygen atoms and hence the structure is linear. i.e*

$$O = C = 0$$

Similarly; CS_2 *molecule is also linear because all the four valence electrons of carbon atom form double bond with sulphur atoms and hence the structure is linear. i.e*

$$S = C = S$$

Q43: **NH_3 *can form coordinate covalent bond with* H^+ *ion. Explain.***

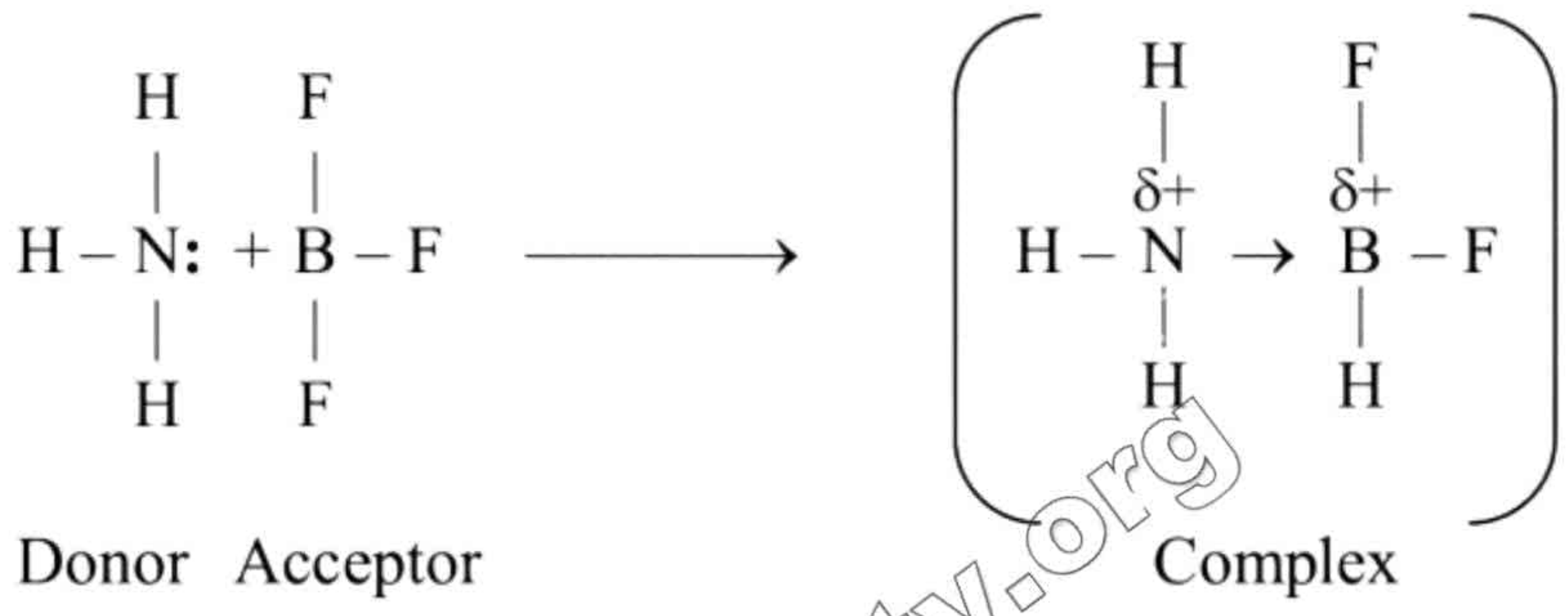
Ans: There is a lone pair in valence shell of N in NH_3*. It can donate to* H^+ *ion to make coordinate covalent bond.*

$$H_3N + H^+ \longrightarrow \underset{\textit{Ammonium ion}}{NH_4^+}$$

Q44: ***What is coordinate covalent bond? Give example.***

Ans: Co-ordinate covalent bond:

A bond which is formed between two atoms in which the shared pair of electrons are donated by one of the bonded atoms. The atoms which donate pair is called donor. The atom which accepts pair is called acceptor.



Q45: ***Why ionic bond is stronger than covalent bond?***

Ans: Ionic compounds are formed by complete transfer of electron between elements. There are strong electrostatic forces of attractions is present among oppositely charged ions and large amount of energy is required to break them.

On the other hand, covalent bond is formed by mutual sharing of electrons between elements. There are weak van der wall's forces are present among atoms and small amount of energy is required to break them

Q46: ***Explain the limitations of Lewis model.***

Ans: Limitations of Lewis model:

Lewis model does not account for,

- *Shapes of molecules.*
- *Various energy transitions as evident by spectroscopic techniques.*
- *Molecular Geometries, bond polarities, bond distances.*

Q47: ***Lone pair of electrons on an atom occupies more space. Give reasons.***

Ans: The electron pairs of lone pairs occupy more space than the bond pairs.

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.

Q48: ***In*** NH_3 ***the bond angle Is 107.3° but in*** NF_3 ***the bond angle is 102°. Explain.***

Ans: NH_3:

In NH_3, N *is surrounded by four electron pairs. Three are bond pairs and one is lone pair. The lone pair present on N exerts greater repulsion on bond pair. Hence bond angle reduce from* 109.5° *to* 107. 5°.

When these H−*atoms are substituted by more electronegative atom like* F *in* NF_3 *bond angle is further reduced. It is because of polarity of* N−F *bond. Hence bond angle is reduced to* 102°.

Q49: ***Write order of repulsion between electron pairs according to VSEPR theory.***

Ans: The repulsion between electron pairs decreases in the following order:

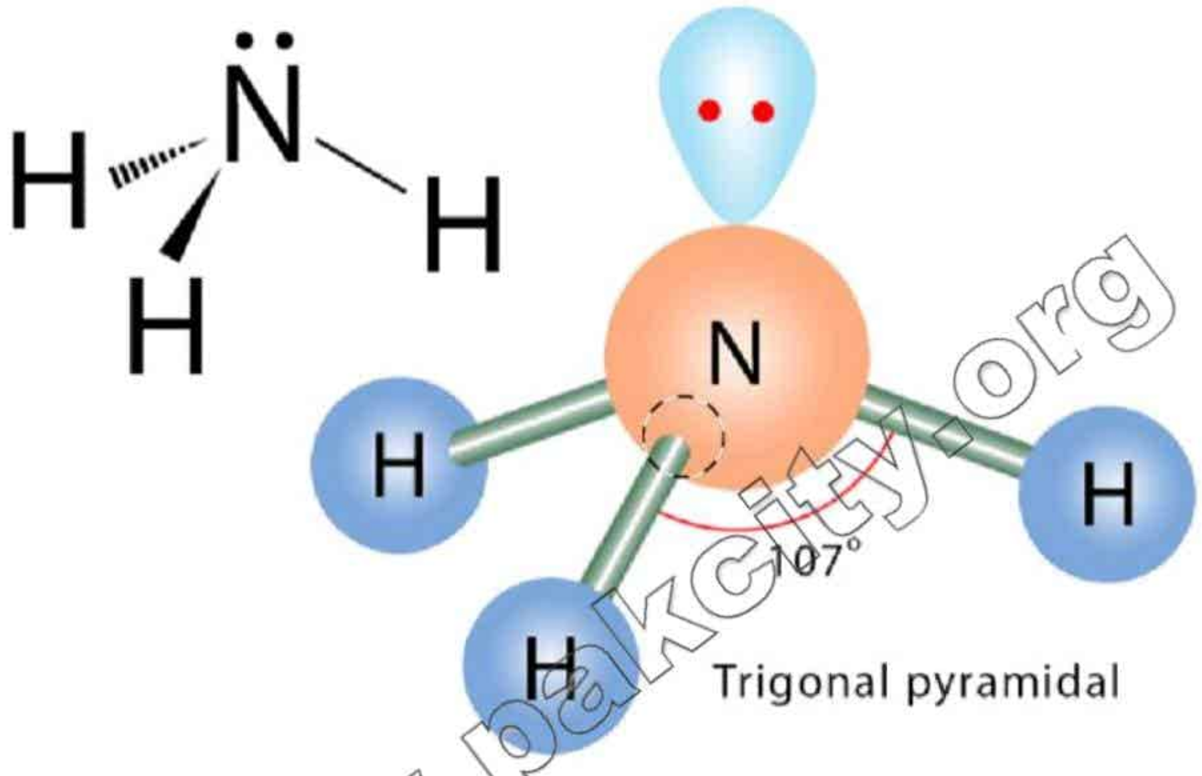
Lone pair − Lone pair > Lone pair − Bond pair > Bond pair − Bond pair

Q50: ***What is the basic assumption for Valence Shell Electron Pair Repulsion (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.

Q51: ***Draw and explain the geometry of*** NH_3 ***molecule according to valence shell electron pair repulsion theory.***

Ans:



In NH_3, N *is surrounded by four electron pair. Three are bond pairs and one lone pair. Due to presence of lone pair tetrahedral geometry changes to trigonal pyramidal with angle* 107.5°.

Q52: ***Define lone pair and bond pair of electrons.***

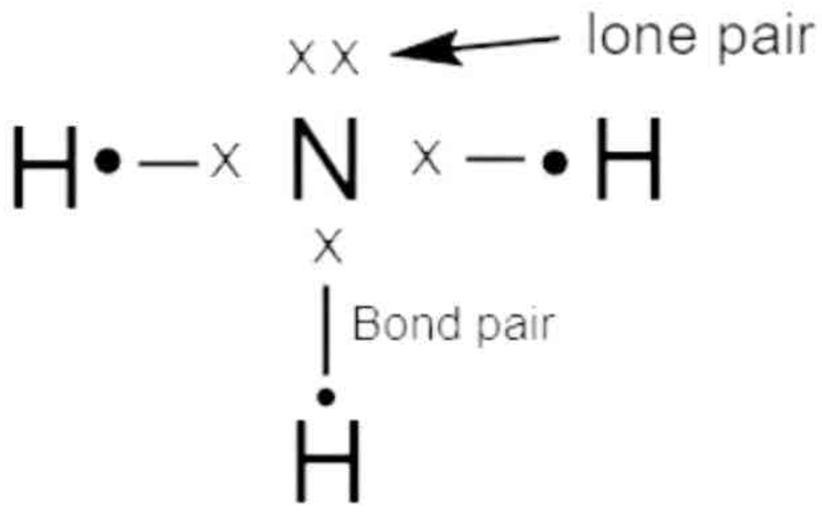
Ans: *Lone pair:*

The electron pair which does not involve in bond formation is called lone pair.

Bond pair:

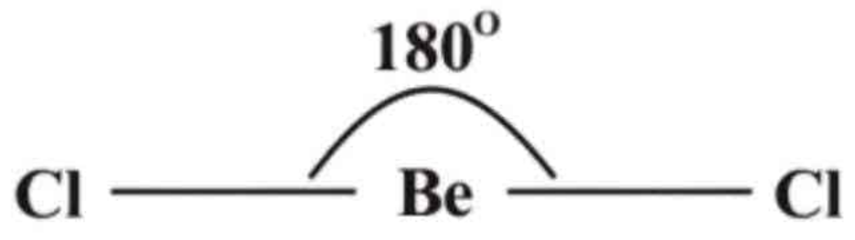
The electron pair which involve in bond formation is called bond pair.

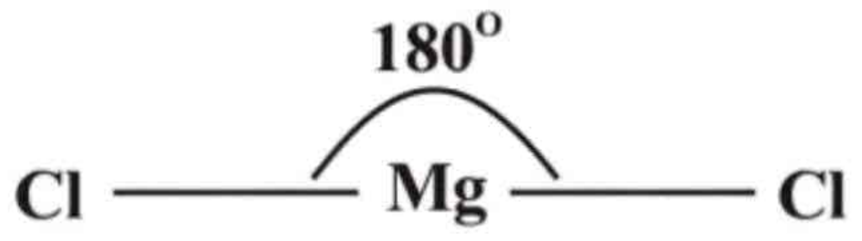
For example:



Q53: ***Discuss*** AB_2 ***type molecules in the light of VSEPR theory. Give two examples.***

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

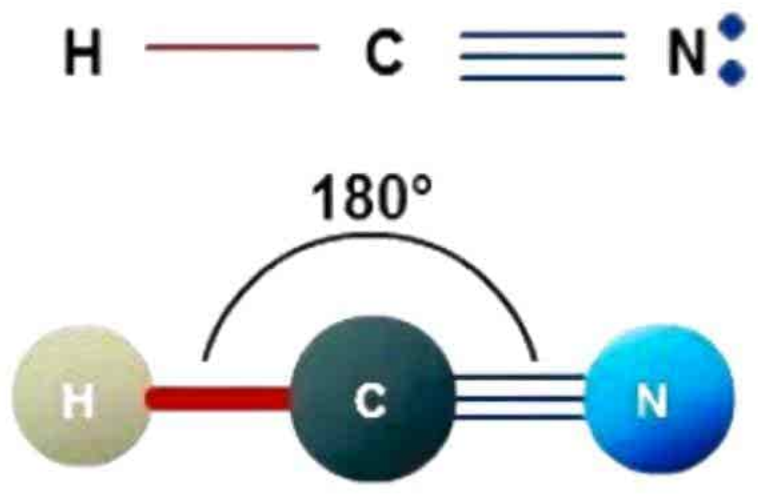




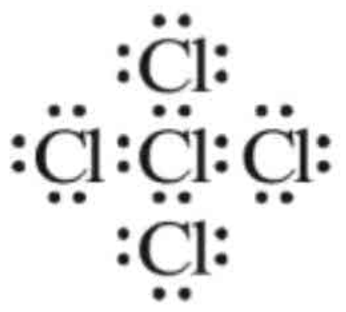
Q54: ***Write the Lewis structures for the following compounds*** **(a) HCN (b) CCl_4**

Ans: *The Lewis structures for the compounds*

(a) HCN



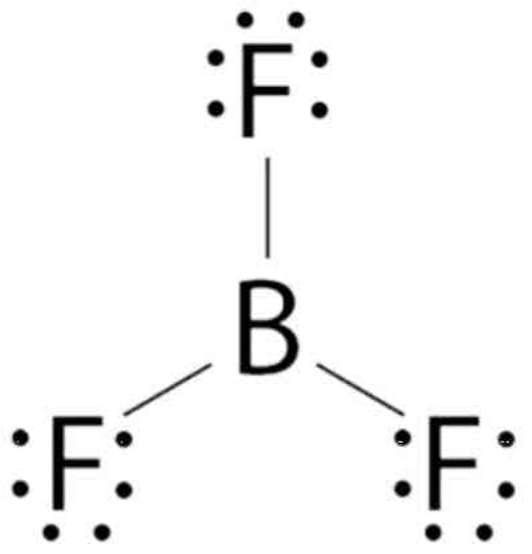
(b) CCl_4



Q55: ***Write the Lewis structures for the following compounds*** **(a) BF_3 (b) CH_4**

Ans: *The Lewis structures for the compounds*

(a) BF_3



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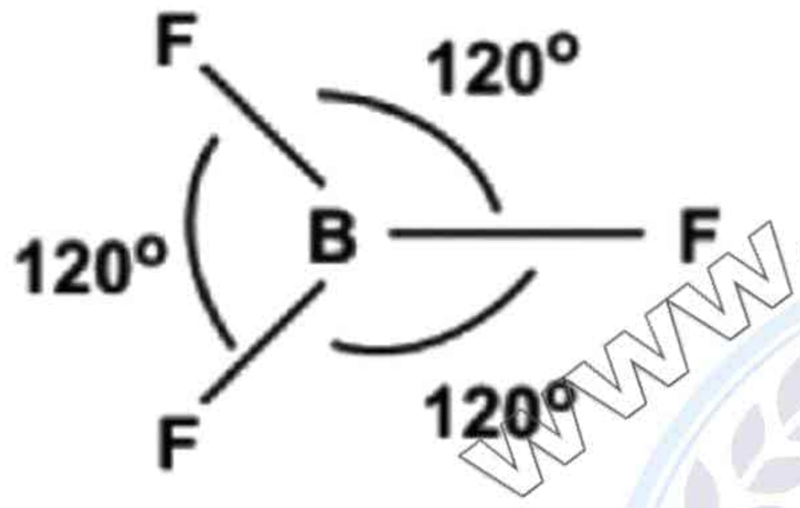
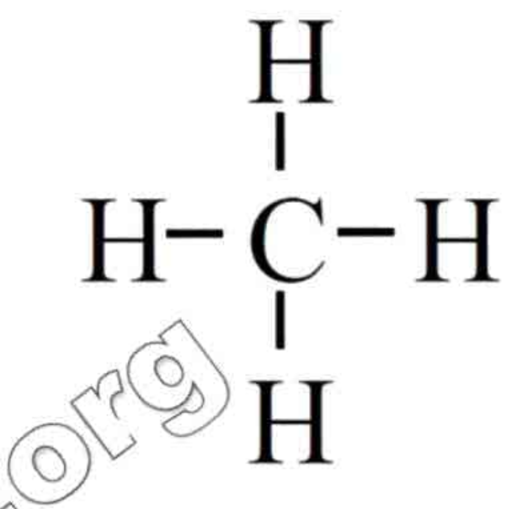
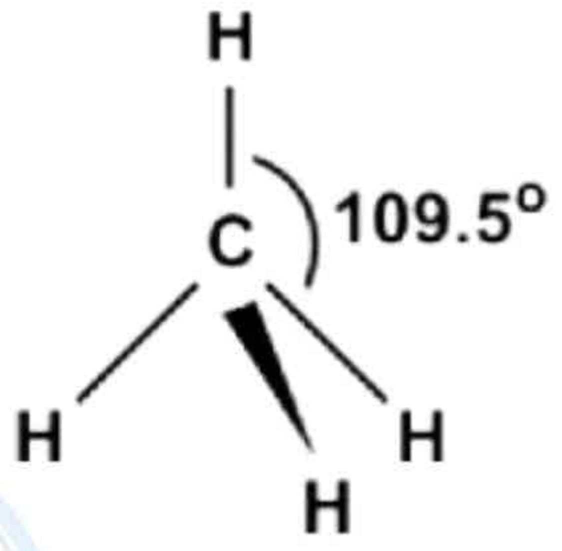


(b) CH_4

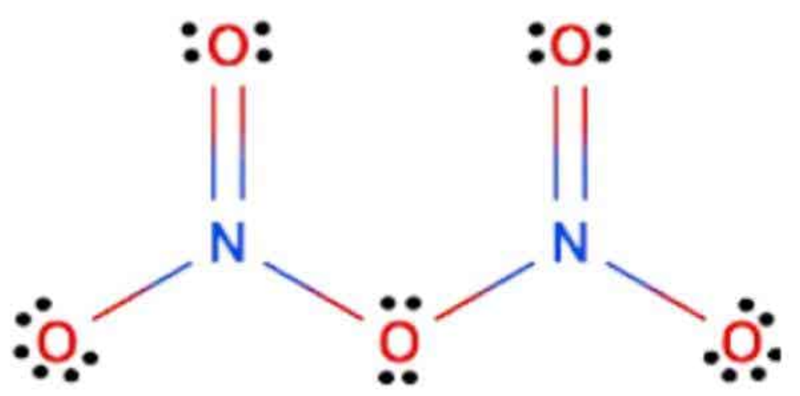




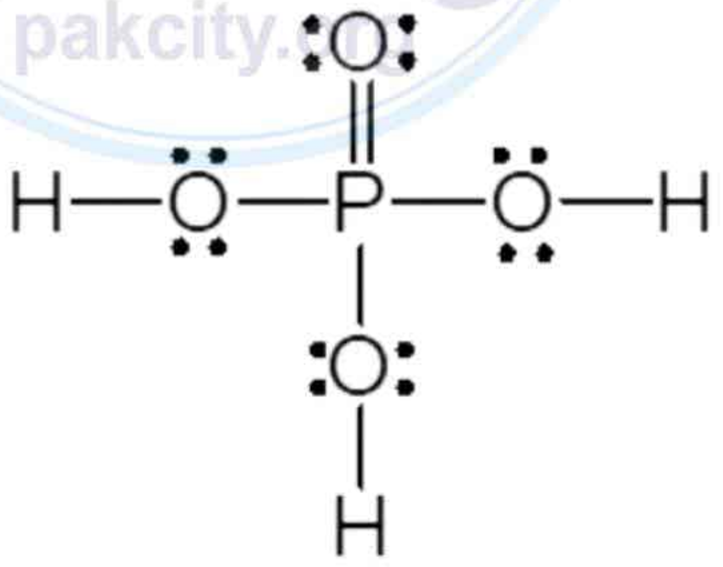
Q56: ***Write the Lewis structures for the following compounds*** **(a) N_3O_5 (b) H_3PO_4**

Ans: *The Lewis structures for the compounds*

(a) N_3O_5



(b) H_3PO_4



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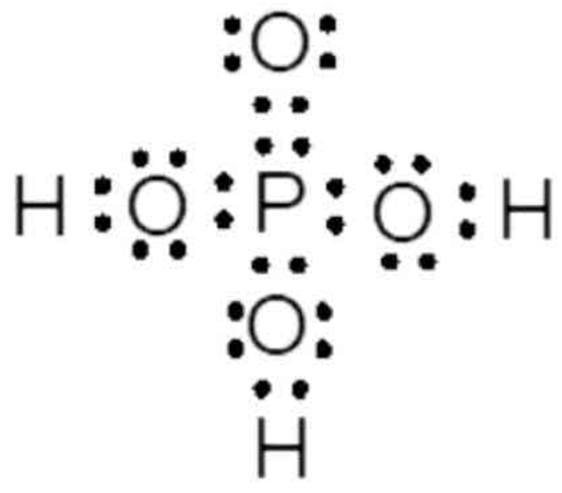
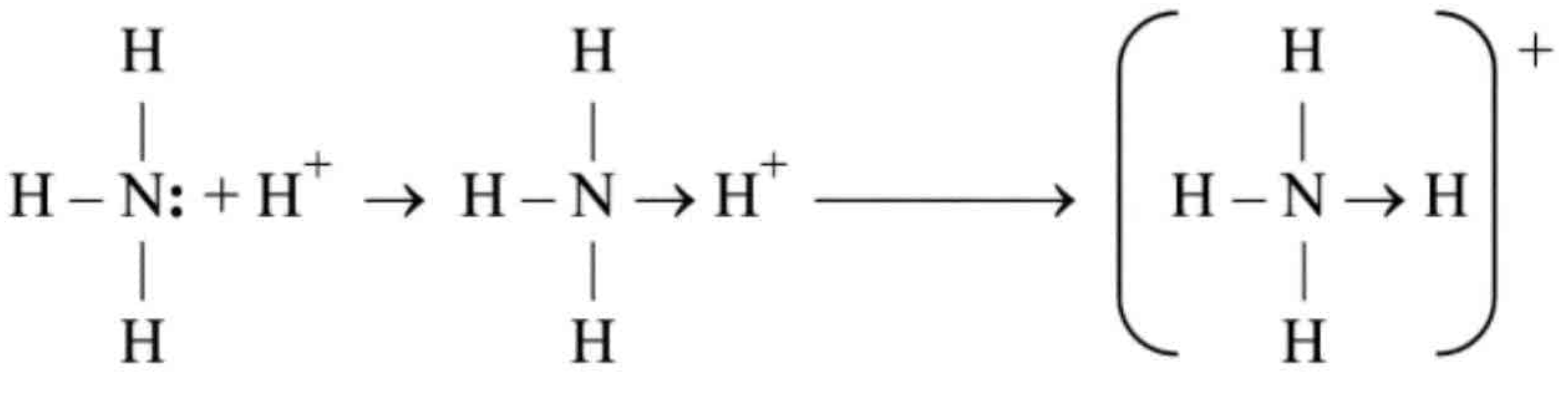


Q57: **NH_3** ***can form coordinate covalent bond with*** **H^+** ***but*** **CH_4** ***cannot. Justify.***

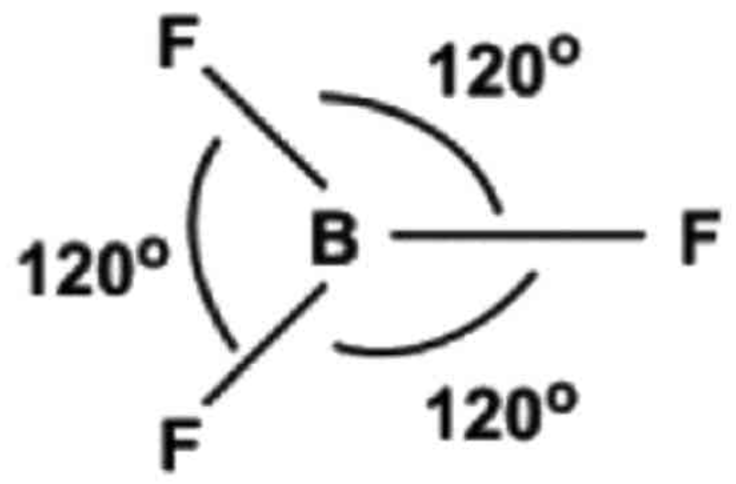
Ans: NH_3 *can form coordinate covalent bond with* H^+ *due to the presence of lone pair in* NH_3 *while* CH_4 *ahas no lone pair and do not form coordinate covalent bond.*



$$H-\underset{H}{\overset{H}{N}}: + H^+ \rightarrow H-\underset{H}{\overset{H}{N}}\rightarrow H^+ \longrightarrow \left[H-\underset{H}{\overset{H}{N}}\rightarrow H\right]^+$$

Q58: ***Why BF_3 molecule is triangular planar.***

Ans: In BE_3 molecule, Boron has three bond pairs in its valence shell which repel each other and adjust at bond angle 120°. So BF_3 molecules have triangular planar geometry.



Q59: ***Explain the structure of NH_3 molecule on the basis of VSEPR theory.***

Ans: According to the VSEPR theory in NH_3 molecule, the cloud of lone pair electrons (non-bonding electrons) spread out more than that of bonding electrons. As a result, somewhat large lone pair charge clouds tend to compress the bond angles in test of the molecules.

$$N_7 = 1s^2, 2s^2, 2p_x{}^{\uparrow}, 2p_y{}^{\uparrow}, 2p_z{}^{\uparrow}$$

The non-bonding electron in 2s orbitals take 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 form 109.5° to 107. 5°. This effect compels ammonia to assume a triangular pyramidal geometry instead of tetrahedral as in methane.

Q60: ***Bond angle in water is 104.5° instead of 109.5° give reason.***

Ans: Water (H_2O) a triatomic molecule is expected to be an AB_2 type linear molecule like $BeCl_2$ and CO_2. But, experimental evidences confirm a bent or angular geometry. VSEPR theory, successfully justifies the experimental results by arguing the participation of lone pairs, in addition to bond pairs in determining overall geometry of water H molecule.

$$O = 1s^2, 2s^2, 2p_x^2, 2p_y{}^{\uparrow}, 2p_z{}^{\uparrow}$$

Two of the corners of a tetrahedron are occupied by each of the two lone pairs and their repulsive action among themselves and on bond pairs, the bond angle in H_2O is further reduced to 104. 5°.

Q61: ***Deviation in bond angle in water H_2O molecule is more than in ammonia NH_3 molecule although both have tetrahedral structures.***

Ans: There are four electron pairs in both NH_3 and H_2O.

In NH_3 there are three bonding electron pairs and one lone pair. Lone pair exerts a repulsion of bonding pair due to which bond angle reduces to 107.5°.

While in case of H_2O there are two bonds pairs and two lone pairs on oxygen atom which exert greater repulsion and hence angle reduce from 109.5° to 104.5°

Q62: ***Give any two postulates of valence shell electron pair repulsion theory.***

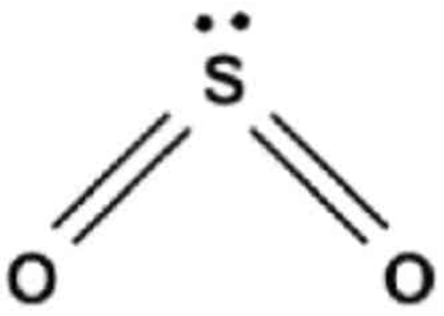
Ans: Valence shell electron pair repulsion theory has following postulates:

- The electron pairs are arranged at; maximum distance apart from each other's to minimize repulsion.
- A lone pair occupied more space than a bond pair.
- Both bond pair and lone pair of electrons are involved in determining the geometry of molecules.

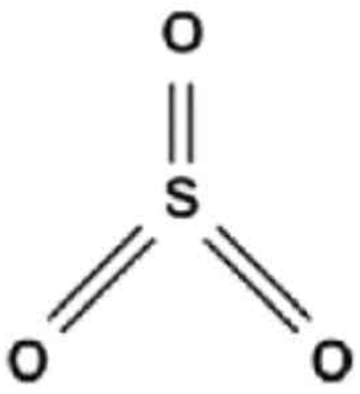
Q63: ***Draw the geometry of*** **SO_2** ***and*** **SO_3** ***on the basis of VSEPR Theory.***

Ans:

SO_2

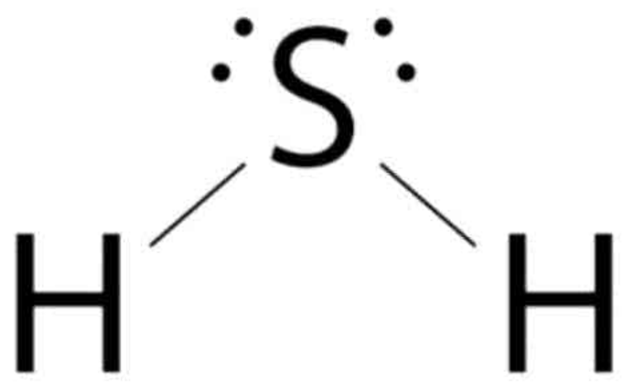


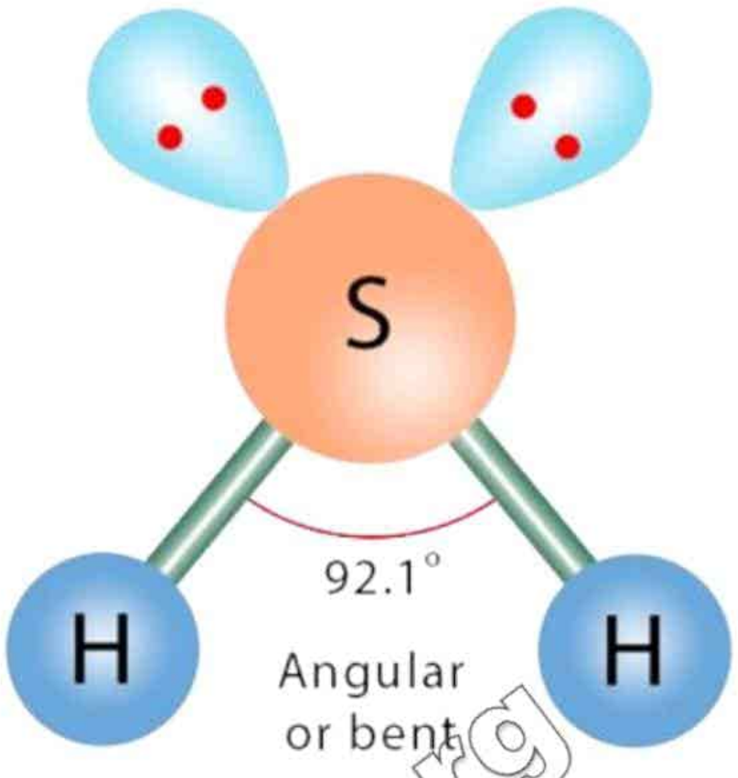
SO_3



Q64: ***Explain geometry of*** **H_2S** ***molecule on the basis of VSEPR Theory.***

Ans: *In valence shell of Sulphur, there are two lone and two bond pairs. They repel each other at maximum distance. As lone pair has greater repulsion with bond pairs, so geometry of* H_2S *is bent with bond angle* 92°.





Q65: ***Both*** **NH_3** ***and*** **BF_3** ***are tetra atomic but different geometries why?***

Ans: *In* NH_3 *Charge cloud of lone pair electrons (non-bonding electrons) spread out more than that of bonding electrons. As a result, somewhat large lone pair cloud tends to compress the bond angles in rest of the molecules. So,* NH_3 *has trigonal pyramidal geometry.*

While in BF_3 *central atom B has three bonding electron pair, which are arranged at maximum distance apart at a mutual angle of* 120° *giving triangular planner geometry.*

Q66: ***Write two points of valence Bond Theory.***

Ans: *Valence Bond Theory has following points:*

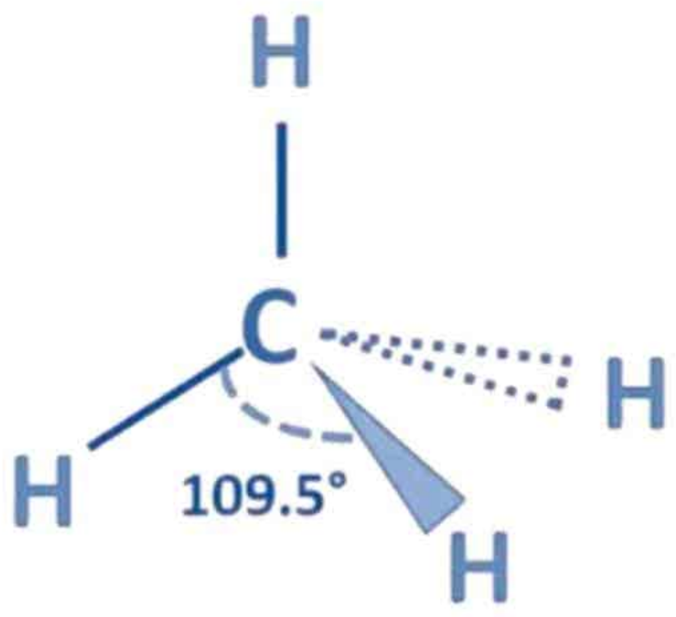
- *Larger the overlap stronger the bond.*
- *The two overlapping, orbitals must be half filled.*
- *The atomic orbitals of valence shell of combining atoms overlap to form covalent bond.*

Q67: ***Sigma Bond is stronger than Pi—Bond. Why?***

Ans: *Strength of bond depends upon the overlapping region. Greater the overlapping, greater will be the strength of bond. Sigma bond has greater overlapping region than Pi—bond. That's why sigma bond is stronger than Pi—bond.*

Q68: ***Tell the number of electron pairs shared in*** **CH_4** ***and how many bonds are polar?***

Ans:

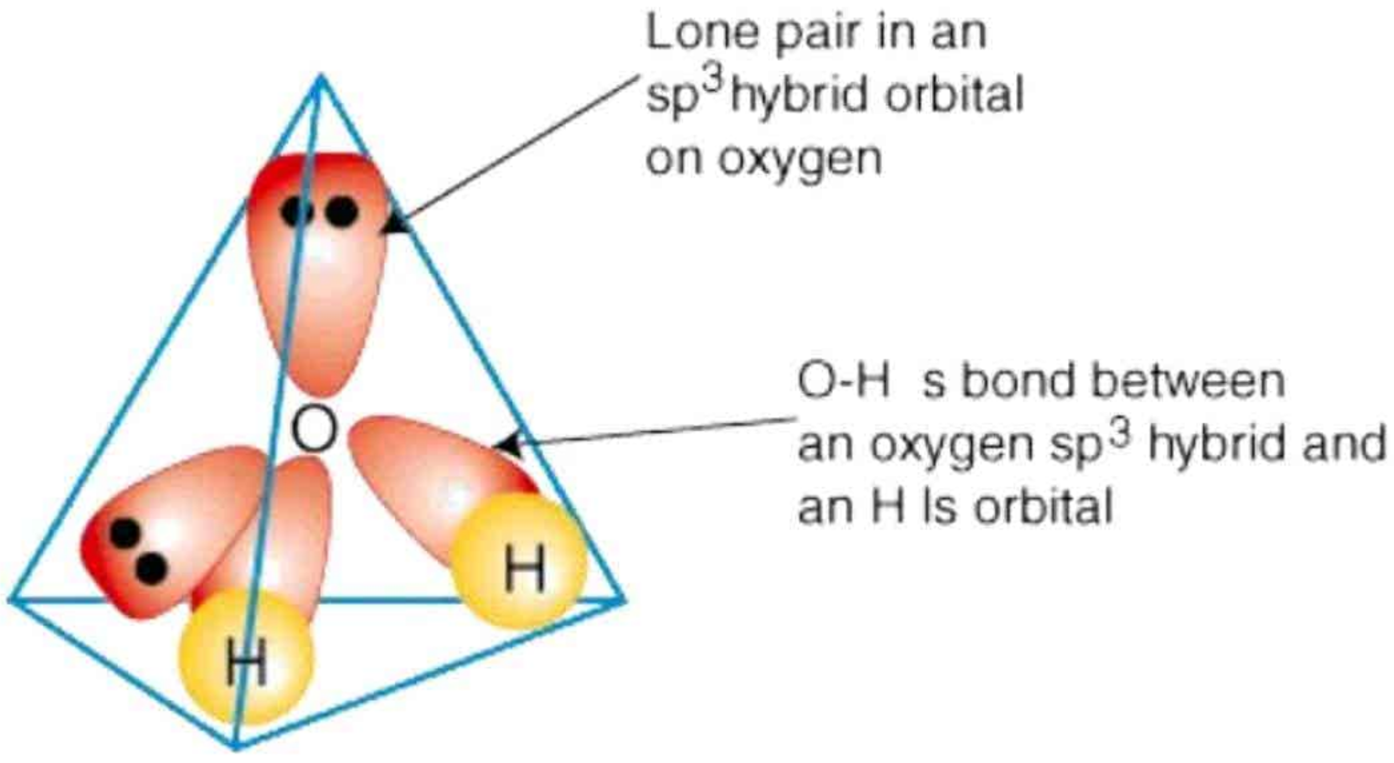


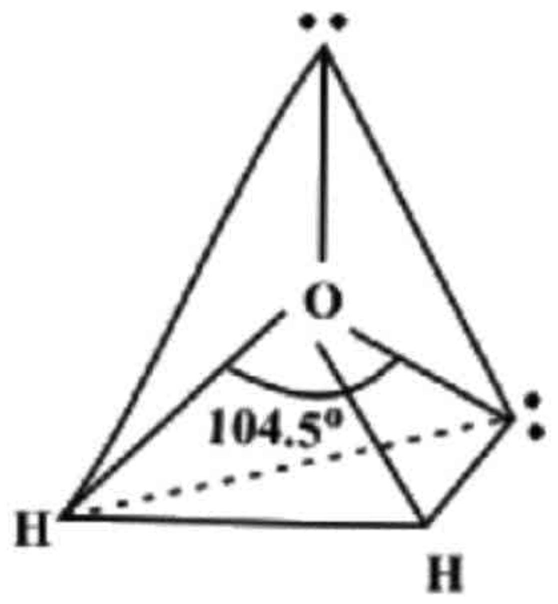
Molecular geometry of CH_4

In this case four electrons paired are shared. All the bonds are polar but total molecule is non-polar due to zero dipole moment. Individual bond moment cancel each other.

Q69: ***Draw the structure of H_2O Molecule on the basis of VBT and explain it.***

Ans:





Here, 2S *and three* 2P *orbitals of oxygen hybridized to form four* SP^3 *hybrid orbitals which will have a tetrahedral arrangement. Two hybrid orbitals are completely filled by the two available lone pair of electrons. The remaining two half—filled hybrid orbitals undergo* $SP^3 - S$ *Overlaps with* H *atoms to form two sigma bonds.*

Q70: ***Define orbital hybridization of orbitals.***

Ans: *Orbital hybridization:*

The process in which atomic orbitals of different energy and shape are mixed together to form a new set of orbitals having same shape and energy is called orbital hybridization.

Q71: ***How bond length is affected by change of hybridization state?***

Ans: *Moreover, hybridization scheme involved, also explains the shortening of bonds due to the predominant participation of* s*—orbital. Since, the* 2s*—orbital of carbon has smaller mean radius than the* 2p*—orbitals, It would be expected that greater thes character in the hybrid orbitals used the shorter will be the bond distance.*

Thus, the C − C *bond lengths are* 154, 133 *and* 120 pm *for ethane, ethene and ethyne respectively where* s *orbital contribution increases from* sp^3 *to* sp*. Further, p—bonding also reduces the internuclear bond distance.*

Q72: ***π —bonds are more diffused than sigma bonds. How?***

Ans: *Sigma bond is formed by head on overlap of atomic orbitals of atoms. Therefore electron density is largely present between the nuclei.*

However Pi—bond is formned by sideways overlap of atomic orbitals. Therefore electron density is largely present above and below the plane of nuclei. Hence Pi—bond.is more diffused than sigma bond.

Q73: ***Define degenerate orbitals. How degeneracy of orbitals can be lifted?***

Ans: *Degenerate orbitals:*

The orbitals which have same energy are called degenerate orbitals. e.g. P subshell has three degenerate orbitals. In presence of magnetic field these are oriented along X, Y, Z axes, and called P_x , P_y , P_z.

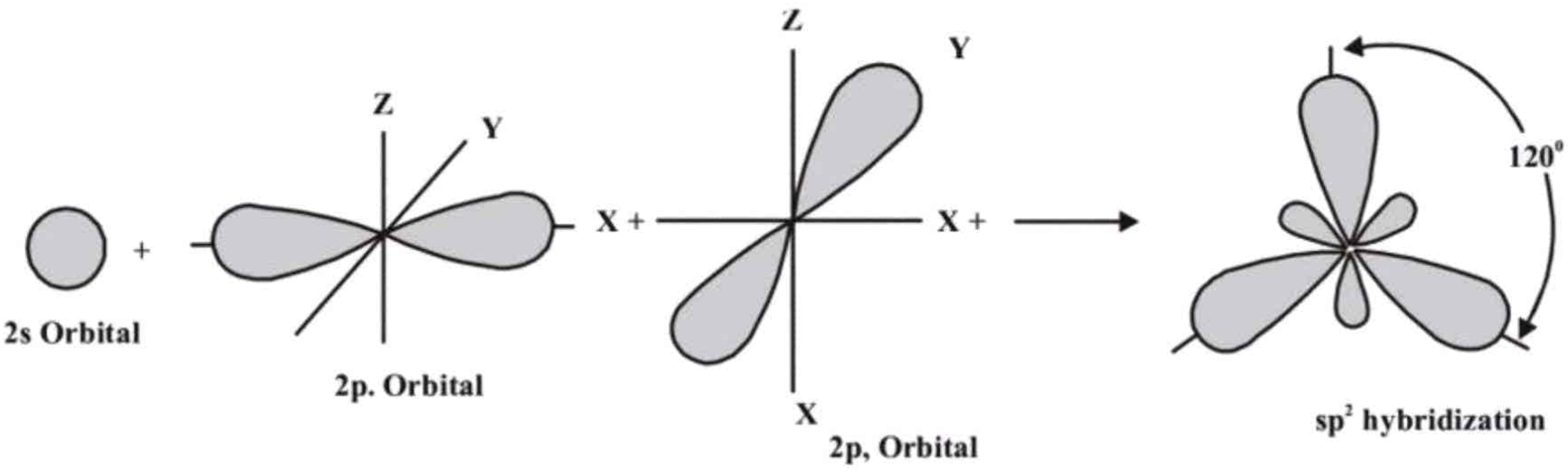
Q74: ***Describe sp^2 hybridization. Mention the molecule with sp^2 hybridization.***

Ans: SP^2*—Hybridization:*

Mixing of one s and two P *orbitals to give three* sp^2 *hybrid orbitals is called* sp^2 *hybridization.*

In BF_3, B *is* sp^2 *hybridized.*

In ethene, C *is* sp^2 *hybridized.*



Three sp^2 hybridized orbitals in one plane and at 120° to each other.

Q75: ***Differentiate between bonding and anti−bonding molecular orbitals.***

Ans: The difference between bonding and anti−bonding molecular orbitals is:

Bonding molecular orbitals	***Anti-bonding molecular orbitals***
The atomic orbitals of bonding atoms overlap to form same number of molecular orbitals. The molecular orbitals which have energy lower than original atomic orbits are called bonding molecular orbitals. i.e, 6, π	*The molecular orbitals which have high energy than original atomic orbitals are called anti-bonding molecular orbitals. i.e,* $6^*, \pi^*$

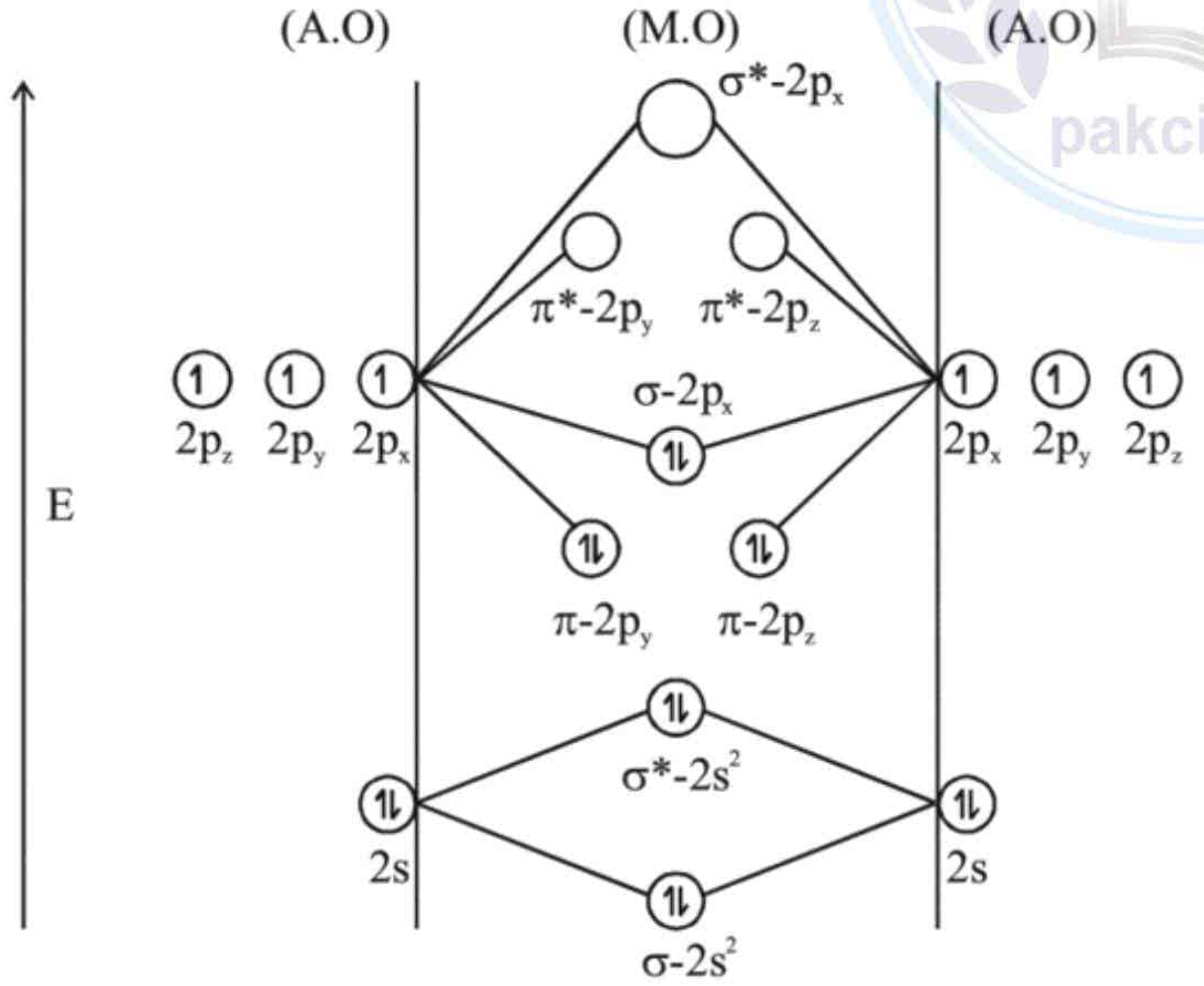
Q76: ***Why the energy of anti−bonding molecular orbital is higher than corresponding bonding molecular orbitals?***

Ans: When atomic orbitals are combining they form molecular orbitals. Some of these molecular orbitals are stable while others are unstable. Those orbitals which are stable have lower in energy than parent orbitals. And those which have higher energy are unstable.

Hence anti-bonding molecular orbitals have higher energy than bonding molecular orbitals.

Q77: ***Draw molecular orbital picture of*** N_2 ***molecule and also calculate its bond order.***

Ans:



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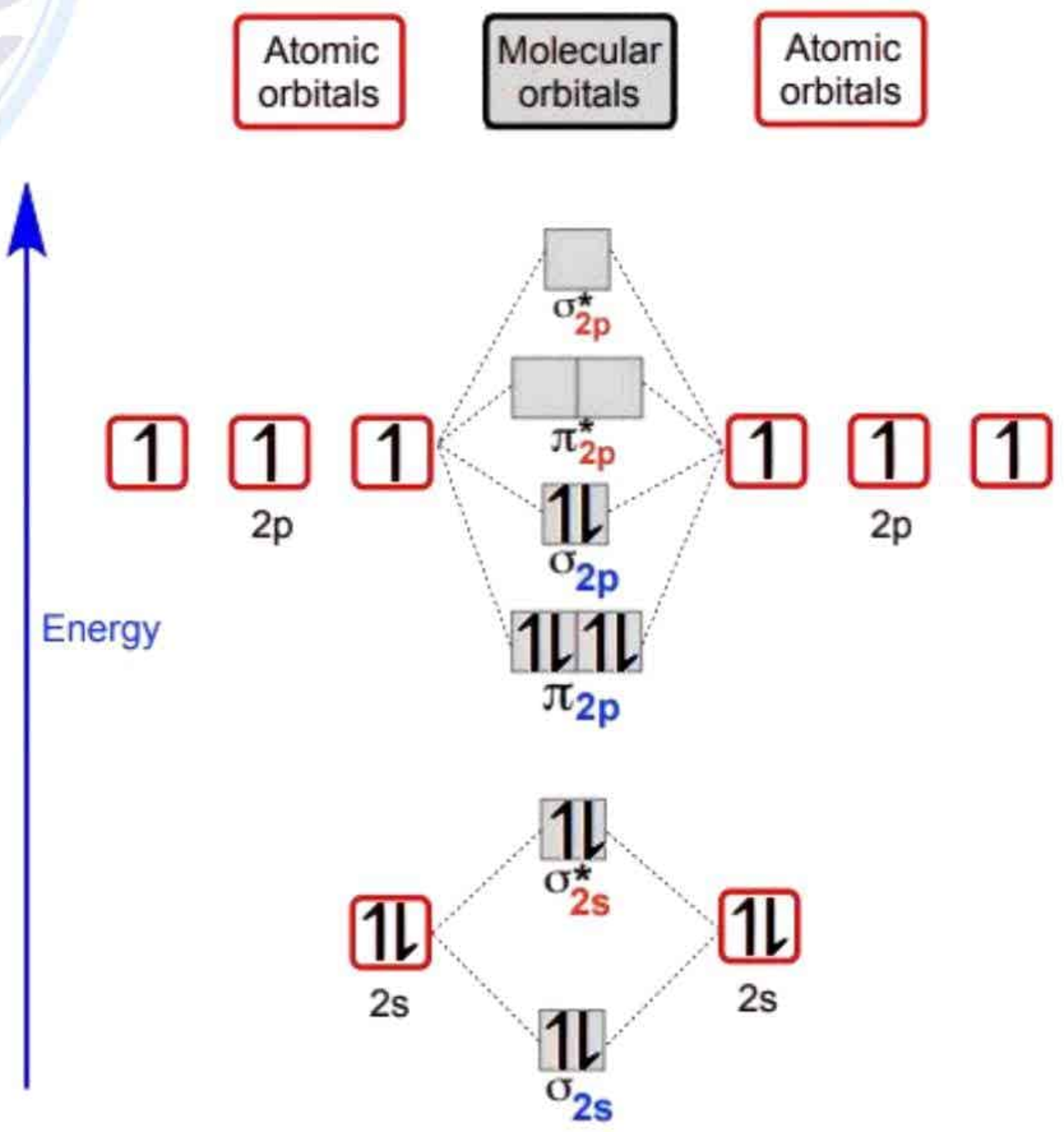


The bond order of N_2 *molecule is* $\frac{6-0}{2} = \frac{6}{2} = 3$ *which correspond to the triple bond consisting of one sigma and two pi bonds.*

Q78: ***What is paramagnetic character?*** ***OR***

Why oxygen molecules show paramagnetic behavior?

Ans: *That property of a substance by which it is attracted towards the magnet due to the presence of unpaired electrons is called paramagnetic character.*

The orbitals accounts admirably for the paramagnetic properties of oxygen. This is one of the greatest successes of the molecular orbital theory. Liquid O_2 *is attracted towards the magnet. Anyhow, when two more electrons are given to* O_2 *it becomes* O_2^{-2}. *The Para magnetism vanished. Similarly in* O_2^{+2} *the unpaired electrons are removed and paramagnetic property is no more there.*

Q79: ***Discuss He-molecules on the basis of MOT.*** ***OR***
Why He_2 ***does not exist under Normal Condition?***

Ans: *The electronic configuration of* He *is* $1s^2$. *Each* He *atom contributes two electrons. Two electrons enter bonding molecular orbital* $\sigma(1s)$ *orbitals and the remaining two electrons go to anti−bondingo* $\sigma^*(1s)$ *molecular orbital.*

The bond order for He_2 *is zero* $\frac{2-2}{2} = \frac{0}{2} = 0$ *and thus* He_2 *molecule is not formed.*

Q80: ***Represent the molecular orbitals of*** N_2 ***molecule in the increasing order of energy.***

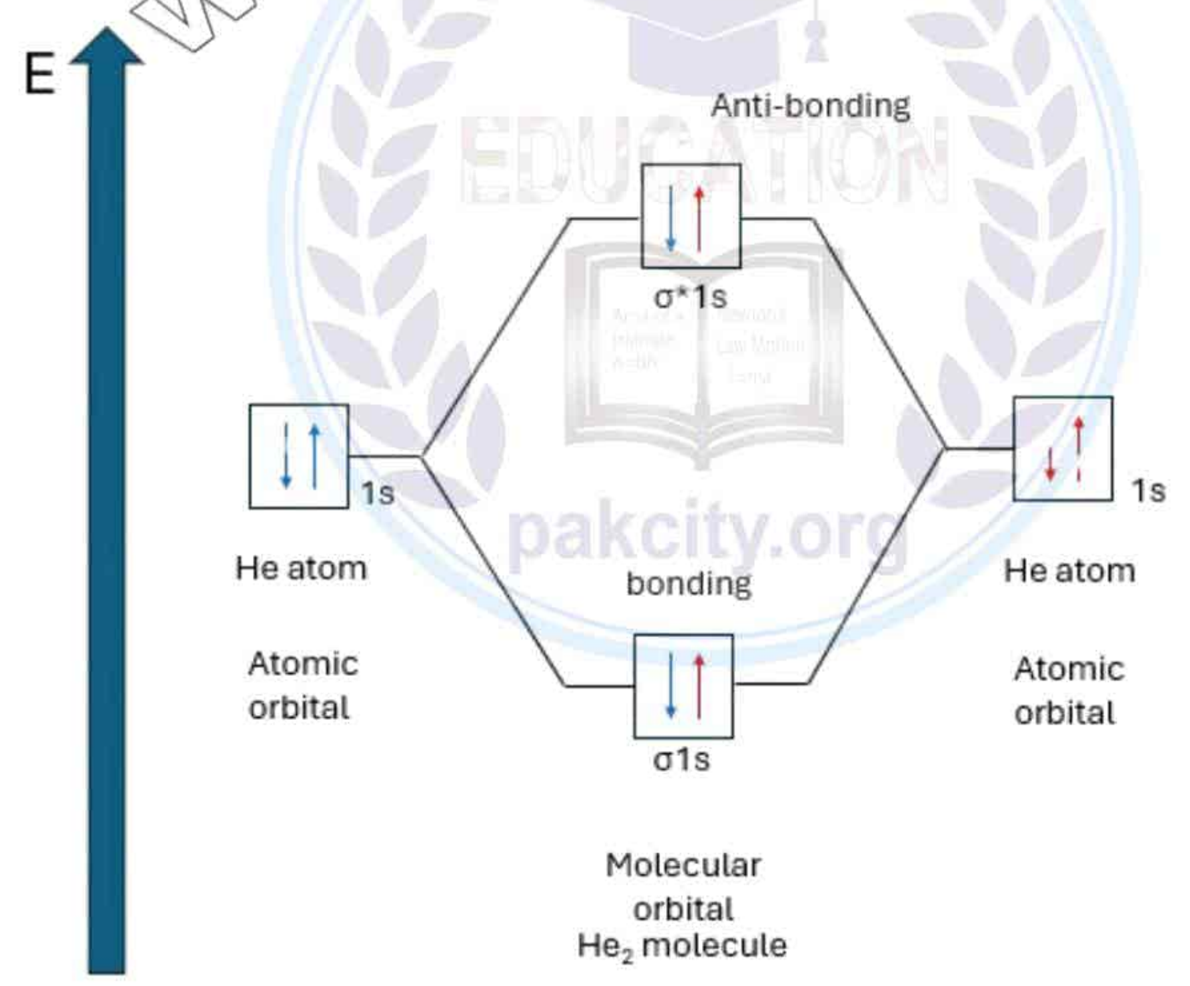
Ans: *The molecular orbitals of* N_2 *in increasing order of their energy are as follows:*

$$\sigma(1s^2) < \sigma^*(1s^2) < \sigma(2s^2) < \sigma^*(2s^2) < \pi(2p_y^2) = \pi(2p_z^2) < \sigma(\ 2p_x^2\ ) < \pi^*\left(2p_y\right) < \sigma^*(2p_x)$$

$$B.O = \frac{6-0}{2} = \frac{6}{2} = 3$$

Q81: ***Draw the hypothetical orbital picture of*** He_2 ***molecule according to M.O.T.***

Ans: *According to MOT when electrons of Helium are distributed two electrons go into bonding molecular orbital and two into anti bonding molecular orbital. Hence all electrons are paired and* He *show diamagnetic behavior.*



Bond order:

$$B.O = \frac{\text{No.of electrons in B.M.O} - \text{No.of electrons A.B.M.O}}{2}$$

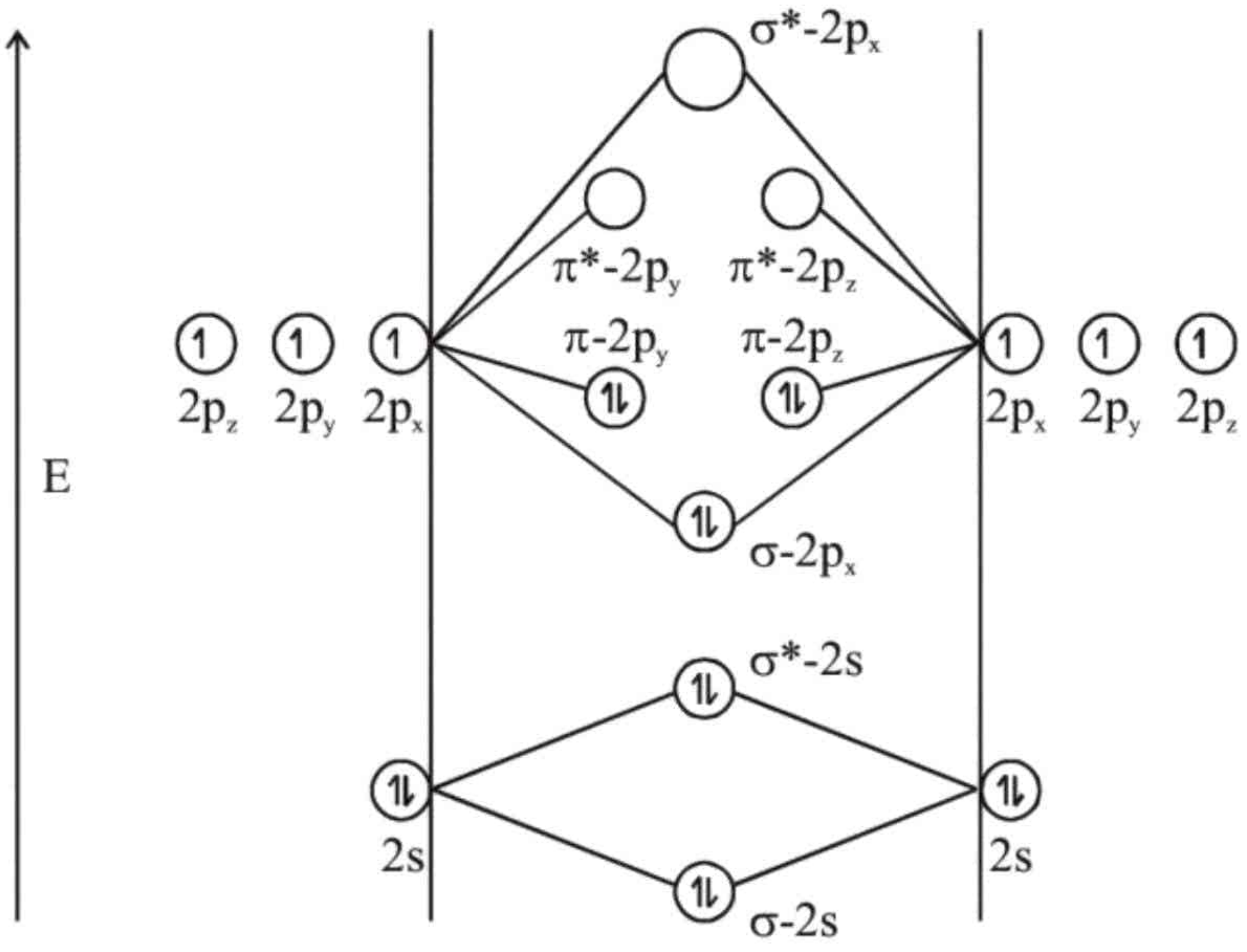
$$B.O = \frac{2-2}{2} = 0$$

He *has zero bond order.*

Q82: ***Define bond order and what bond order is of*** O_2^{+2}.

Ans: *Bond order:*

The number of bonds formed between two atoms by overlap of atom orbitals is called bond order.



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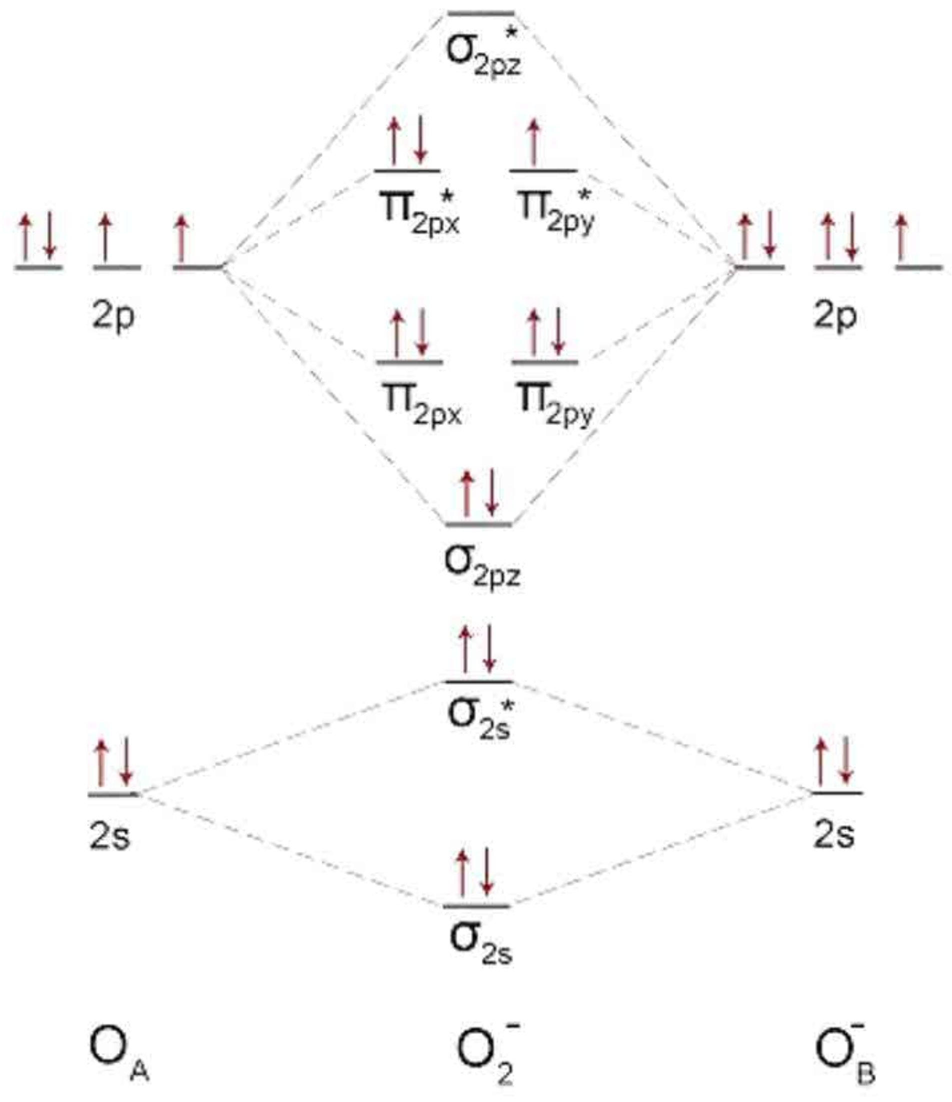


$$\text{Bond order} = \frac{6-4}{2} = 0$$

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

Q83: ***Differentiate between atomic and molecular orbital.***

Ans: The difference between atomic and molecular orbital is:

Atomic orbital	***Molecular orbital***
It is the region around the nucleus of an atom in which the probability for finding the electrons is maximum. It has no sharp boundaries and can have different shapes.	*The orbital which has more than one nuclei buried in its electronic cloud is called molecular orbital.*

Q84: ***MOT is Superior to VBT explain.***

Ans: MOT is Superior to VBT is:

MOT	***VBT***
➢ *According to this theory atomic orbitals combine to give new orbital called each other to form covalent bond.* ➢ *The overlapping orbitals lose their identity.* ➢ *MOT can well explain paramagnetic and diamagnetic behavior.*	➢ *According to this theory valence shell atomic orbitals of atoms overlap with combine to give new orbital called molecular orbitals.* ➢ *The overlapping orbitals retain their individual identity.* ➢ *V.B.T fails to explain paramagnetic and diamagnetic behaviors.*

Q85: ***Bond distance is compromise distance between two atoms. Explain with reason.***

Ans: When the atoms approach each other to form a bond they attract each other and the energy of the system is lowered. When they reach at a certain distance their forces of attractions are maximum. On further coming close they start repelling each other and thus the energy of the system increases.

Now the two atoms try to remain at a compromising distance whereas the energy of the system is minimum. This bond distance is known as compromising distance between two atoms.

Q86: ***The abnormality of bond length and bond strength in*** **HI** ***is less prominent than that of*** **HCl*****. Why?***

Ans: *Greater the electronegativity difference between bonded atoms shorter the bond length and stronger the bond.*

$$\begin{array}{ll} H-Cl & H-I \\ E.N\ 2.1\quad 3.00 & 2.1\quad 2.5 \\ 3.00-2.1=0.9 & 2.5-2.1=0.4 \end{array}$$

Since the difference in electro negativities is greater in $H-Cl$*. Therefore abnormality of bond length and bond strength is greater than that of* $H-I$ *where the difference in electronegativity value is very small.*

Q87: ***How the criteria of electronegativity help us to understand the nature of bond?***

Ans: *The criteria of electronegativity help us to understand the nature of bond is:*

E.N difference b/w bonded atoms	***Nature of bond***
➢ *Zero*	➢ *Non-polar covalent bond*
➢ *More than 1.7*	➢ *Ionic bond*
➢ *Less than 1.7*	➢ *Polar covalent bond*
➢ *Equal to 1.7*	➢ *50% ionic and 50% covalent*

Q88: ***What is difference between bond length and bond energy?***

Ans: *The difference between bond length and bond energy is:*

Bond length	***Bond energy***
The average distance between the nuclei of two atoms forming a covalent bond is called bond length. *Example: In* $C-C$ *bond length is* 154 pm.	*The amount of energy required to break all bonds of particular type in one mole of a substance is called bond energy.*

Q89: ***Why*** CH_4 ***molecule is called tetrahedral? Mention maximum number of bond angles in it?***

Ans: *The hybrid orbitals are oriented in space in such a manner by the overlap of* sp^3 *hybrid orbitals. The hybrid orbital of carbon with* 1s *orbital of four hydrogen atoms separately to form four sigma bonds. The molecule, thus formed, possesses a tetrahedral geometry.*

The four $C-H$ *bonds which result from* sp^3-s *overlaps are directed towards the corners of a regular tetrahedron. There are six bonds angles each* 109.5° *the tetrahedral structure of* CH_4 *has four faces, four corners and six edges.*

Q90: ***Define Covalent Radius with an example.***

Ans: *Covalent radius:*

Covalent radius is defined as half of the single bond length between two similar atoms covalently bonded in a molecule. The covalent radius of hydrogen, for example is 37.7 pm.

It is half of the single bond length (75.4 pm) *between the two* H *atoms in* $H-H$ *molecule.*

Q91: ***Why sigma bond is stronger than Pi−bond?***

Ans: *Strength of bond depends upon the overlapping region. Greater will be overlapping region, greater will be the strength of bond. Sigma bond has greater overlapping region than Pi−bond, so it is stronger than Pi−bond.*

Q92: ***Differentiate between Sigma and Pi−bond.***

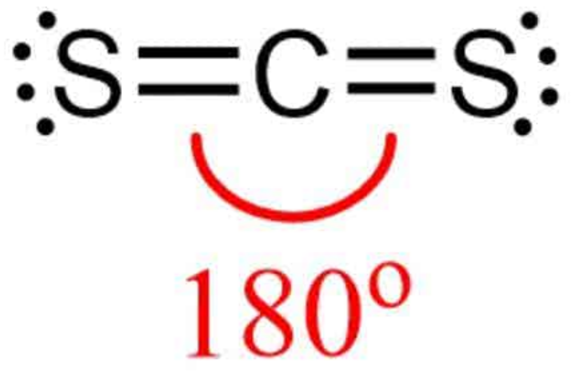
Ans: *The difference between Sigma and Pi−bond is:*

Sigma bond (σ)	***Pi−bond*** (π)

When partially filled atomic orbitals overlap in linear way and probability of finding electron is on the line joining two nuclei, sigma bond is formed.	*When partially filled atomic orbitals overlap in parallel fashion, the probability of finding electron lies above and below the line joining two nuclei π −bond is formed.*

Q93: ***Why CS_2 molecule is linear while SO_2 is angular?***

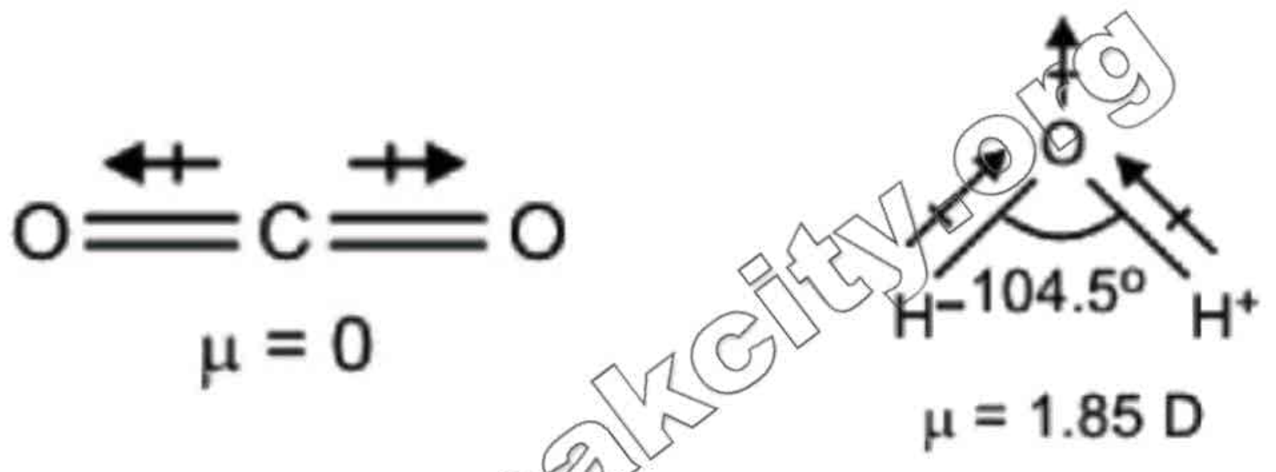
Ans: CS_2 *molecule is linear because all four valence electrons of* C *atom form double bond with sulphur atoms and hence structure is linear.*



In case of SO_2 *molecule, central 'S' has atom two bond pairs and one lone pair due to lone pair repulsion, angle is less than* 120° *and structure is angular.*

Q94: ***CO_2 and H_2O both are triatomic molecules but the dipole moment of CO_2 is zero and that of H_2O is 1.85 D. Give reason.***

Ans: *In case of* CO_2 *the individual bond moment cancel each other due to linear structure of* CO_2 *so, it has zero dipole moment.*



While in H_2O *individual bond moment do not cancel each other. Due to angular structure it has some dipole moment value.*

Q95: ***What information about the molecular structure are obtained from the dipole moment?***

Ans: *Diploe moment help to determine the shape of molecules.*

Diatomic molecule:

These molecules are always linear structure.

Triatomic molecules:

If triatomic molecules have $\mu = 0$ *then has linear structure. It molecule has some value of dipole moment then structure is bent. Dipole moment also tells about Tetra-atomic and pent atomic molecules.*

Q96: ***Why BF_3 is non−polar but SO_2 is polar?***

Ans: BF_3 *is a non-polar molecule because it has symmetrical triangular planner geometry having zero dipole moment. This is all due to the cancellation of individual bond moments.*

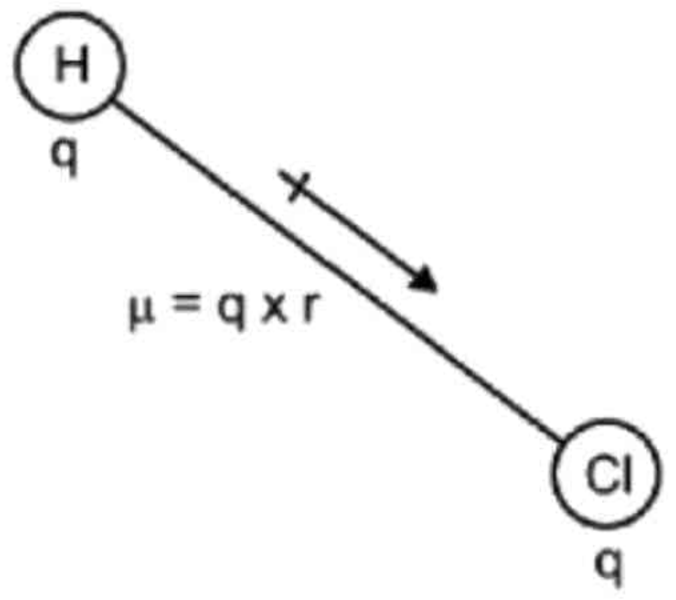
Whereas triatomic molecule SO_2 *is bent like* H_2O*. Dipole moment of* SO_2 *is* 1.61 D *which ruled out its linear structure.*

Thus due to dipole moment present in SO_2 *molecule, it is polar.*

Q97: ***Define "Dipole Moment" and give its mathematical form and S.I unit.***

Ans: *Dipole moment:*

The product of the magnitude of charge on both end and the distance between the positive and negative centers is called dipole moment.



Dipole moment and its vector representation

Mathematical form:

If 'q' *is magnitude of charges and* 'r' *is the distance between them, then*

$$\mu = q \times r$$

Units:

S.I unit of dipole moment is mC. *(Meter Coulomb)*

If two charges are separated by 100 pm then

$$\mu = (1.602 \times 10^{-19}\text{C}) \times (100 \times 10^{-12}\text{m})$$

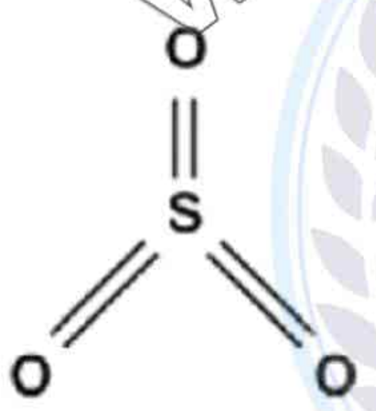
$$\mu = 1.602 \times 10^{-29}\ \text{mC}$$

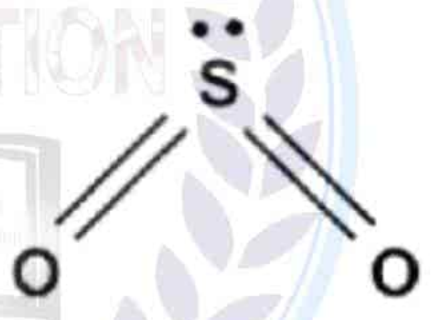
Q98: ***Why the dipole moment of*** **SO_2 *is* 1.61 D** ***but that of*** **SO_3** ***is zero?***

Ans: *In* SO_2 *both* $S = O$ *bonds are polar.* SO_2 *molecule has an angular structure. This structure is not symmetrical. In this case individual bond moments do not cancel each other's.*

Therefore it has dipole moment. But in case of SO_3 *this is a symmetrical molecule in which individual bond moment cancels the effect of each other's, so it has zero dipole moment.*

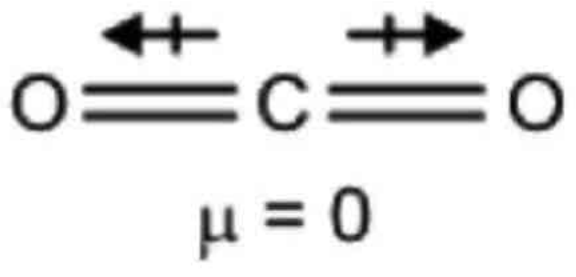
Structure is shown below:

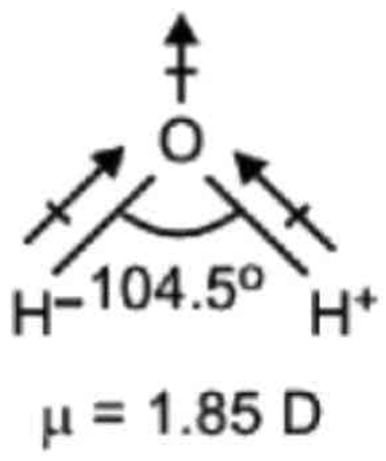




Q99: **CO_2 *molecules have linear structure while* H_2O *have angular structure. Comment.***

Ans: *In case of* CO_2 *the value of dipole moment is zero which suggests linear structure. The individual bond moment cancel each other.*



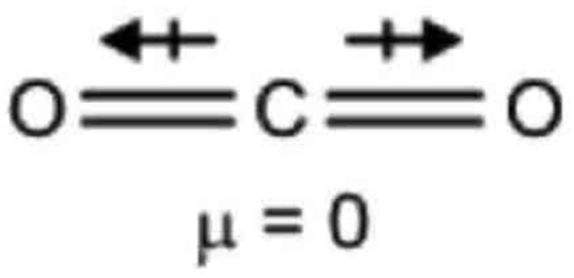


In this case of H_2O *the individual bond moment do not cancel each other which suggest angular structure for them.*

Q100: ***Separate the following molecules as polar or non−polar.***

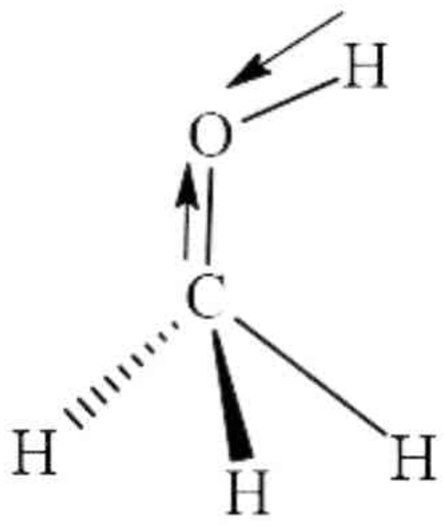
(a) CO_2 (b) CH_3OH (c) CCl_4 (d) HF

Ans: (a) CO_2:



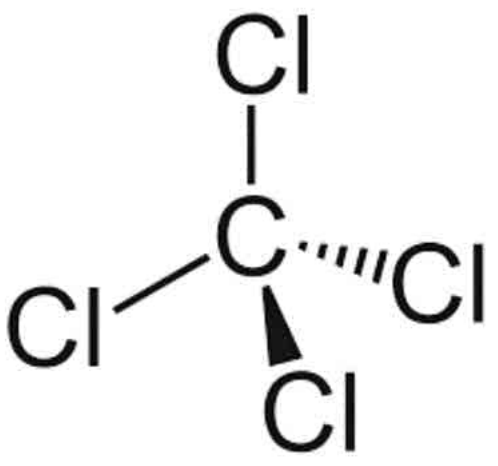
In this case $C = O$ *bond is a polar but individual bond moment cancel each other's so overall molecule has* $\mu = 0$ *so,* CO_2 *is non−polar.*

(b) CH_3OH:



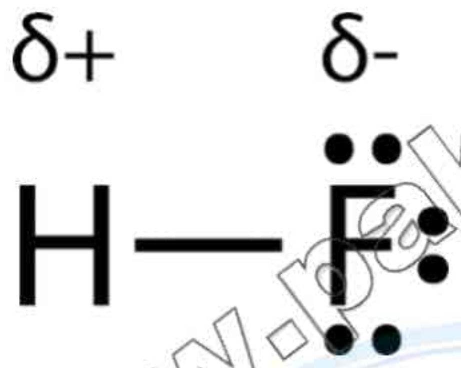
In this case individual bond moments do not cancel each other. Hence it is a polar molecule.

(c) CCl_4:



This is a symmetrical tetrahedral molecule in which individual bond moment cancel each other so overall it is non−polar.

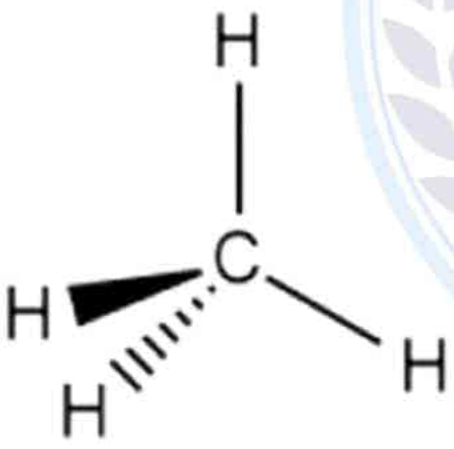
(d) HF:



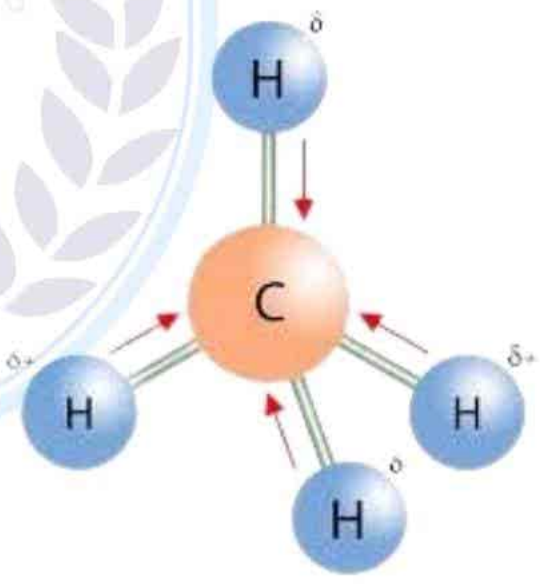
It has some definite dipole moment value so it is a polar molecule.

Q101: ***Why the dipole moment of*** CH_4 ***is zero?***

Ans:



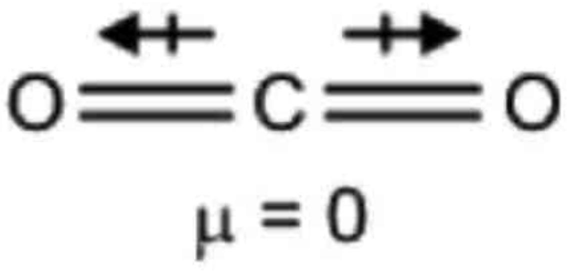
OR

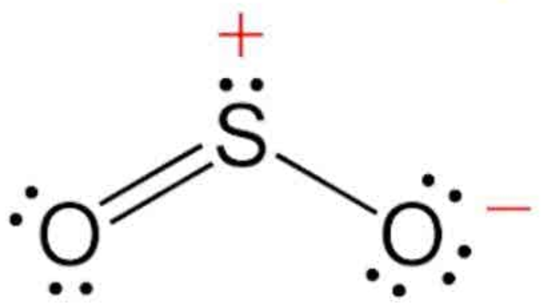


In this case individual bond moments cancel each other that are why CH_4 *has zero dipole moment.*

Q102: ***The dipole moment of*** CO_2 ***is zero but*** SO_2 ***is not zero. Why?***

Ans:

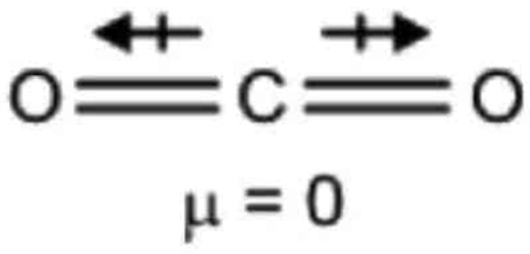


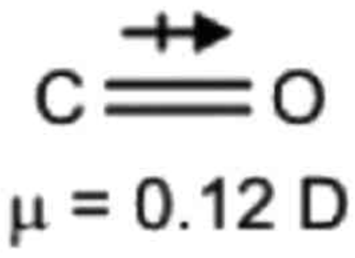


In this case individual bonds cancel each other and its dipole moment is zero. $\mu = 1.64$ D SO_2 *has angular structure because it has some value of dipole moment.*

Q103: ***The dipole moment of*** CO_2 ***is zero but that of*** CO ***is*** **0.12 D*****. Give reason.*** ***OR***
Why CO ***is polar and*** CO_2 ***is non-polar?***

Ans: CO_2 *has zero dipoles and its geometry is linear. Polarity of molecule is cancelled out. In* CO *single dipole is present which is not cancelled out as shown below:*





Q104: ***If charge of any one of poles of a dipole is*** $\mathbf{1.6022 \times 10^{-19}C}$ ***and two poles are separated by*** $\mathbf{1A^\circ}$ ***then calculate dipole moment in unit of Debye.***

Ans: Let us consider a hypothetical molecule $(A^+ - B^-)$, *or a unit negative charge separated from a unit positive charge by distance* $r = 100\ pm\ (1\mathring{A})$. *The dipole moment of such a molecule can be calculated by multiplying the distance* 100pm *to charge of one electron or proton is* $1.6022 \times 10^{-19}C$.

$$\mu = (1.6022 \times 10^{-19}C) \times (100 \times 10^{-12}m)$$

$$\mu = 1.6022 \times 10^{-29}\ mC$$

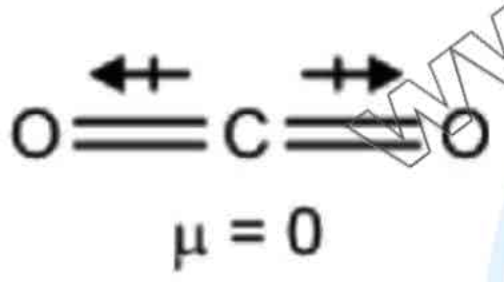
Q105: ***Why no bond can have 100% ionic character?***

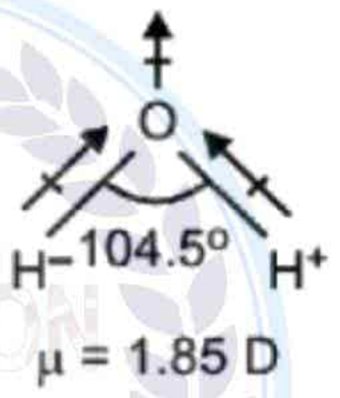
Ans: Elements of group IA *an* IIA *when react with elements of group* VIA *and* VIIA *they form ionic bond. In ionic bond the cations and anions are considered as spheres.*

Due to attraction of positive and negative ions, electron could distorted and attraction toward each other. In this way covalent characters are also present in ionic bond.

Q106: ***Why*** $\mathbf{CO_2}$ ***has zero dipole moment and that of water has value of dipole moment 1.85 D?***

Ans: CO_2 *has zero dipole moment because* CO_2 *has a linear structure, where the dipoles being equal and opposite, cancel out each other's effect.*





The dipole moment of water is 1.85 D *which rules out its linear structure. The calculation shows that water has an angular structure with a bond angle* 104.5 *between the two* O − H *bonds. A linear* H_2O *molecule* (H − O − H) *would have zero dipole moment.*

Q107: ***Why the melting points, boiling points, heat of sublimation and heat of vaporization of electrovalent compounds are higher as compared with those of covalent compounds?***

Ans: Melting and boiling points of ionic or electrovalent compounds are usually high. 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 founds in solids, liquids and gases. Covalent compounds have less attractive force between atoms.

Q108: ***Why covalent compounds react slowly than the ionic compounds?***

Ans: Reactions of covalent compounds are generally slow. Because no electrical forces are present in them to speed up the reaction and energy is require to break old bonds and to form new bonds.

Q109: ***Reactions between ionic compounds are very rapid, give reason.***

Ans: The ionic compounds exist in the form of ions in an aqueous solution. The chemical reaction between ions occurs rapidly.

For example, addition of silver nitrate solution to sodium chloride solution produces a white precipitate of silver chloride instantaneously. The reaction is rapid because on mixing the solutions, no bonds have to be broken, only a new bond is formed. The ionic compounds have already been broken while forming their aqueous solutions.

Q110: ***What is the effect of bonding on solubility of a compound?***

Ans: Mostly, ionic compounds are soluble in water but insoluble in non−aqueous solvents.

When a crystal of an ionic substance is placed in water, the polar water molecules detach the cations and anions from the crystal lattice by their electrostatic attraction.

In general, covalent compounds dissolve easily in nonpolar organic solvents (benzene, ether, etc.) Here, the attractive forces of solvent molecules are enough for overcoming the intermolecular forces of attraction. Mostly, covalent compounds are insoluble in water. However, some of them dissolve in water due to hydrogen bonding.

Q111: ***Ionic compounds are mostly soluble in water but insoluble in non−polar solvents. Give reason.***

Ans: Ionic compounds are water soluble due to ion−dipole forces. The anions are attracted by partial positive hydrogen of water and cations by partial negative oxygen's.

No such interactions are developed by non−polar solvents.

Chapter : 06 Chemical Bonding

Long Questions

Q1: Explain important points of molecular orbit theory and draw structure of nitrogen (N_2) molecule according to this theory.

Q2: Define dipole moment. Give its SI units. Also write its significance. Explain the geometry of CO_2, BF_3 and H_2O by dipole moment.

Q3: Explain atomic orbital hybridization with reference to structures for C_2H_4 and C_2H_2.

Q4: Explain type of hybridization in H_2O and NH_3.

Q5: Define bond energy. Discuss relation between ionic character and bond energy.

Q6: What is ionization energy and discuss variation of ionization energy in periodic table. Also explain how ionization energy is an index of metallic character? (V.Imp)

Q7: Define electron affinity. Name the factors affecting it. How does it vary in the periodic table? (V.Imp)

Q8: Differentiate clearly between covalent and coordinate covalent bond with the help of two examples in each case.

Q9: Give important postulates of VSEPR theory. (V.Imp) OR

Write the main postulates of VSEPR theory and explain the structure of ammonia on the basis of this theory.

Q10: *Draw the shapes of following molecules according to VSEPR theory.*

(i) NH_3 (ii) BF_3 (iii) H_2O (iv) $BeCl_2$

Q11: *Write down the main postulates of VSEPR theory and discuss the structure of* NH_3 *with reference of this theory.*

Q12: *Explain valance bond theory (VBT) with examples showing the overlapping of different orbitals.*

Q13: *How will you describe paramagnetic character of* O_2 *molecules on basis of MOT?*

Q14: *Write a note on* Sp *hybridization and give example of ethyne.*

Q15: *Explain molecular orbital theory. Give molecular orbital configuration & structure of* O_2.

Q16: *What is atomic orbital hybridization? Explain bonding and structure of ethane* (C_2H_4) *on the basis of hybridization?*

Q17: *Explain* Sp^3 *hybridization with the help of two examples. (V.Imp)* *OR*

Define Sp^3 *hybridization. Draw the structure of ammonia molecule according to hybridization concept.*

Q18: *Define* Sp^2 *hybridization and on its basis explain the structure of ethane. (V.Imp)*

Q19: *Define hybridization of atomic orbitals. Discuss* Sp^3 *hybridization in detail. Draw the structure of methane.*

Q20: *Calculate the bond order of* O_2 *molecule by making energy level diagram. Also show that it is paramagnetic. (V.Imp)*