Multiple Choice Questions

Single Correct Answer Type

Q1. Isostructural species are those which have the same shape and hybridization. Among the given species identify the isostructural pairs.
(a) [NF_3 and BF_3]
(b) [BF^-_4 and NH^+_4]
(c) [BCl_3 and $BrCl_3$]
(d) [NH_3 and NO^-_3]

Sol: (b) NF_3 is pyramidal whereas BF_3 is planar triangular.
BF^-_4 and NH^+_4 ions both are tetrahedral.
BCl_3 is triangular planar whereas $BrCl_3$ is pyramidal.
NH_3 is pyramidal whereas NO_3 is triangular planar.
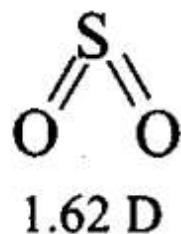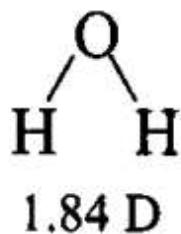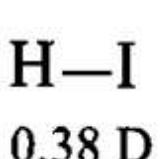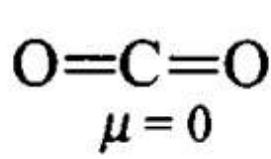
Q2. Polarity in a molecule and hence the dipole moment depends primarily on electronegativity of the constituent atoms and shape of a molecule. Which of the following has the highest dipole moment?
(a) CO_2
(b) HI
(c) H_2O
(d) SO_2

Sol: (c) H_2O will have highest dipole moment due to maximum difference in electronegativity of H and O atoms.

$$O=C=O \qquad H-I \qquad \overset{O}{H \quad H} \qquad \overset{S}{O \quad O}$$
$$\mu = 0 \qquad 0.38\ D \qquad 1.84\ D \qquad 1.62\ D$$

Q3. The types of hybrid orbitals of nitrogen in NO_2, NO_3 and NH_4 respectively are expected to be
(a) sp, sp^3 and sp^2
(b) sp, sp^2 and sp^3
(c) sp^2, sp and sp^3
(d) sp^2, sp^3 and sp

Sol: (b) The number of orbitals involved in hybridization can be determined by the application of formula:

$$H = \frac{1}{2}[V + M - C + A]$$
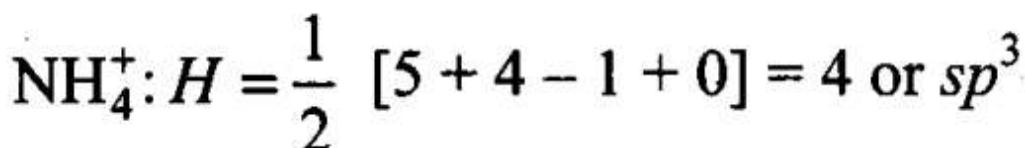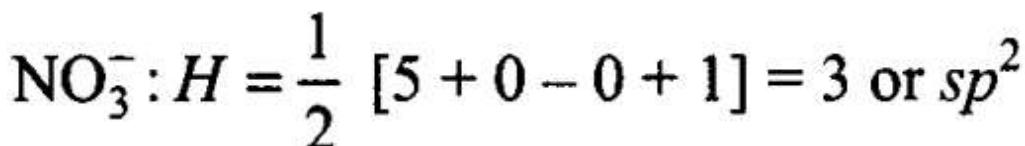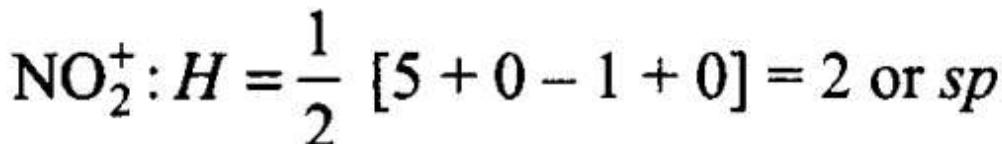
where H = number of orbitals involved in hybridization

V= valence electrons of central atom

M= number of monovalent atoms linked with central atom

C = charge on the cation

A = charge on the anion

$$NO_2^+ : H = \frac{1}{2}[5 + 0 - 1 + 0] = 2 \text{ or } sp$$

$$NO_3^- : H = \frac{1}{2}[5 + 0 - 0 + 1] = 3 \text{ or } sp^2$$

$$NH_4^+ : H = \frac{1}{2}[5 + 4 - 1 + 0] = 4 \text{ or } sp^3$$

Q4. Hydrogen bonds are formed in many compounds e.g., H_2O, HF, NH_3. The boiling point of such compounds depends to a large extent on the strength of hydrogen bond and the number of hydrogen bonds: The correct decreasing order of the boiling points of above compounds is

(a) HF>H_2O>NH_3

(b) H_2O>HF>NH_3

(c) NH_3>HF>H_2O

(d) NH_3>H_2O>HF

Sol: (b) Strength of H-bonding depends on the electronegativity of the atom which follows the order: F > O > N .

Strength of H-bond is in the order:

H....... F > H........ O > H........ N
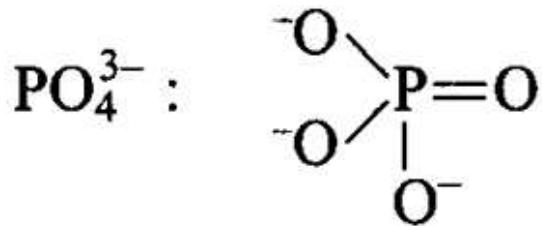
But each H_2O molecule is linked to 4 other H_2O molecules through H-bonds whereas each HF molecule is linked only to two other HF molecules.

Hence, correct decreasing order of the boiling points is H_zO > HF > NH_3.

Q5. In PO_4^{3-} ion the formal charge on the oxygen atom of P – O bond is

(a) +1   (b) -1   (c) -0.75   (d) +0.75

Sol: (b) Formal charge of the atom in the molecule or ion = (Number of valence electrons in free atom) – (Number of lone pair electrons + 1/2 Number of bonding electrons)

$$PO_4^{3-} :$$

$$\text{For charge on oxygen} = 6 - \left(6 + \frac{1}{2} \times 2\right) = 6 - 7 = -1$$

Q6. In NO_3^- ion, the number of bond pairs and lone pairs of electrons on nitrogen atom are

(a) 2, 2   (b) 3, 1   (c) 1,3   (d) 4, 0
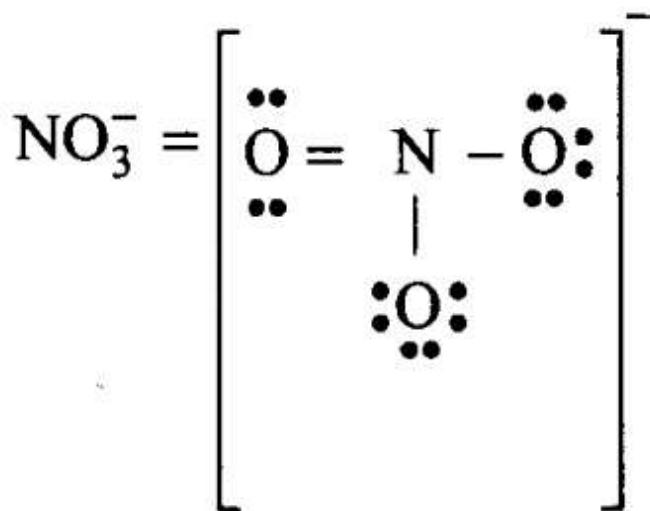
Sol: (d) In N-atom, number of valence electrons = 5

Due to the presence of one negative charge, number of valence electrons = 5 + 1 = 6. One O-atom forms two bonds (= bond) and two O-atom are shared with two electrons of N-atom.

Thus, 3 O-atoms are shared with 8 electrons of N-atom.

Number of bond pairs (or shared pairs) = 4
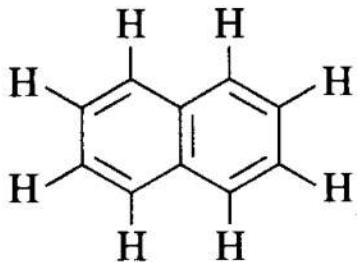
Number of lone pairs = 0

Q7. Which of the following species has tetrahedral geometry?

Sol. (a) $BH_4^- =$ tetrahedral, $NH_2^- =$ V-shape, $CO_3^{2-} =$ triangular planar, $H_3O^+ =$ pyramidal.

8. Number of π bonds and σ bonds in the following structure is
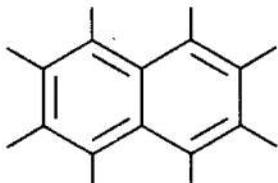
    (a) 6, 19   (b) 4, 20   (c) 5, 19   (d) 5, 20

Sol. (c)

$8C - H$, $11\ C - C\ \sigma$ bond or $19\ \sigma$ bonds and $5\ \pi$ bonds.

Q9. Which molecule/ion out of the following does not contain unpaired electrons?
(a) N^+_2
(b) O_2
(c) O_2^{2-}
(d) B_2

Sol. (c) $N_2^+ = KK\ \sigma 2s^2\ \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2\ \sigma 2p_z^1$

$O_2 = KK\ \sigma 2s^2\ \sigma^* 2s^2\ \sigma 2p_z^2\ \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^1 = \pi^* 2p_y^1$

$O_2^{2-} = KK\ \sigma 2s^2\ \sigma^* 2s^2\ \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^2 = \pi^* 2p_y^2$

$B_2 = KK\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_x^1 = \pi 2p_y^1$

Thus, O_2^{2-} has no unpaired electrons.

Q10. In which of the following molecule/ion all the bonds are not equal?
(a) XeF_4
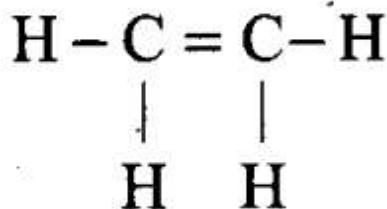
(b) BF^-_4

(c) C_2H_4

(d) SiF_4

Sol: (c) C_2H_4 has one double bond and four single bonds. Bond length of double bond (C = C) is smaller than single bond (C – H).

$$H-C = C-H$$

with H and H below the two carbons connected by single bonds.

Q11. In which of the following substances will hydrogen bond be strongest?
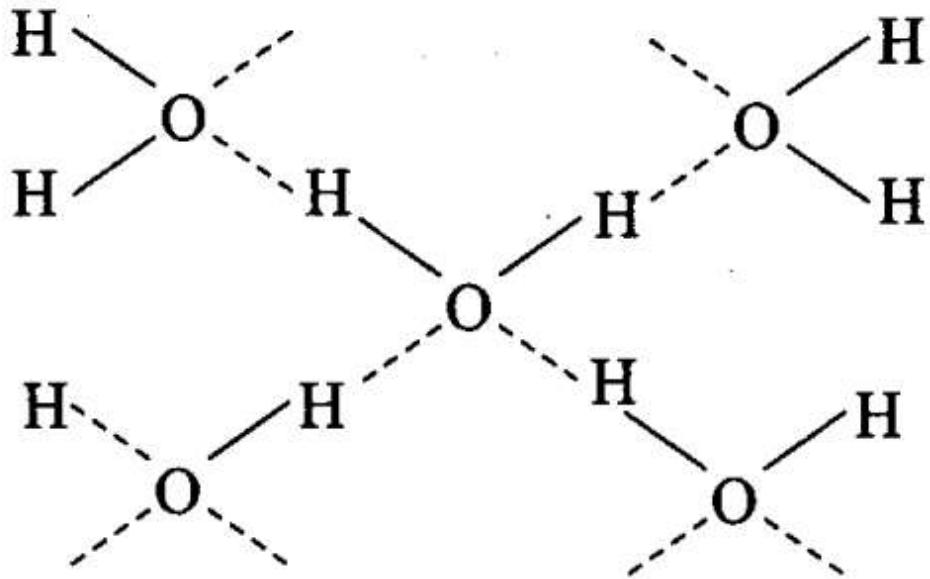
(a) HCl

(b) H_2O

(c) HI

(d) H_2S

Sol: (b) HC1, HI and H_2S do not from H-bonds. Only H_2O forms hydrogen bonds. One H_2O molecule forms four H-bonding.

Q12. If the electronic configuration of an element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$, the four electrons involved in chemical bond formation will be .

(a) $3p^6$

(b $3p^6,\ 4s^2$

(c) $3p^6, 3d^2$

(d) $3d^1, 4s^2$

Sol:(d) In transition elements $(n-1)d$ and ns orbitals take part in bond formation.

Q13. Which of the following angle corresponds to sp^2 hybridisation?
(a) 90° (b) 120° (c) 180° (d) 109°

Sol: (b) sp^2 hybridisation gives three sp^2 hybrid orbitals which are planar triangular forming an angle of 120° with each other.

The electronic configurations of three elements A, B and C are given below.

Answer the questions from 14 to 17 on the basis of these configurations.

A $1s^2 2s^2 2p^6$

B $1s^2 2s^2 2p^6 3s^2 3p^3$

C $1s^2 2s^2 2p^6 3s^2 3p^s$

Q14. Stable form of A may be represented by the formula
(a) A

(b) A_2

(c) A_3

(d) A_4

Sol: (a) The given electronic configuration shows that A represents noble gas because the octet is complete. A is neon which has 10 atomic number.

Q15. Stable form of C may be represented by the formula
(a) C

(b) C_2

(c) C_3

(d) C_4

Sol: (b) The electronic configuration of C represent chlorine. Its stable form is dichlorine (Cl_2), i.e., C_2.

Q16. The molecular formula of the compound formed from B and C will be
(a) BC

(b) B_2C

(c) BC_2

(d) BC_3

Sol: (d) B represent P and C represents Cl. The stable compound is PCl_3 i.e., BC_3.

Q17. The bond between B and C will be

(a) ionic

(b) covalent

(c) hydrogen

(d) coordinate

Sol: (b) Both B and C are non-metals and therefore, bond formed between them will be covalent.

Q18. Which of the following order of energies of molecular orbitals of N_2 is correct?

(a) $(\pi 2p_y) < (\sigma 2p_z) < (\pi^*2p_x) \approx (\pi^*2p_y)$

(b) $(\pi 2p_y) > (\sigma 2p_z) > (\pi^*2p_x) \approx (\pi^*2p_y)$

(c) $(\pi 2p_y) < (\sigma 2p_z) > (\pi^*2p_x) \approx (\pi^*2p_y)$

(d) $(\pi 2p_y) > (\sigma 2p_z) < (\pi^*2p_x) \approx (\pi^*2p_y)$
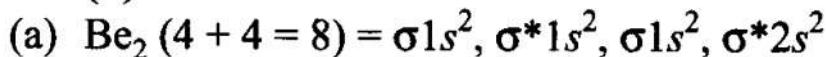
Sol. (a) For N_2, energy of $\sigma 2p_z$ is higher than $\pi 2p_x$ and $\pi 2p_y$. Hence, the order will be $\pi 2p_y < \sigma 2p_z < \pi^*2p_x = \pi^*2p_y$

Q19. Which of the following statements is not correct from the view point of molecular orbital theory?

(a) Be_2 is not a stable molecule.

(b) He_2 is not stable but He^+_2 is expected to exist.

(c) Bond strength of N_2 is maximum amongst the homonuclear diatomic molecules belonging to the second period.

(d) The order of energies of molecular orbitals in N_2 molecule is $\sigma 2s < \sigma*2s < \sigma 2p_z < (\pi 2p_x = \pi 2p_y) < \pi*2p_x = \pi*2p_y) < \sigma*2p_z$.

Sol. (d)

(a) Be_2 $(4+4=8) = \sigma 1s^2, \sigma*1s^2, \sigma 1s^2, \sigma*2s^2$
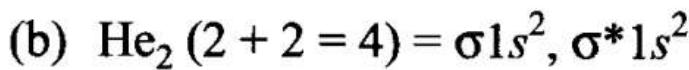
Bond order (BO) $= \dfrac{1}{2}$ [Number of bonding electron (N_b) – Number of anti-bonding electrons N_a]
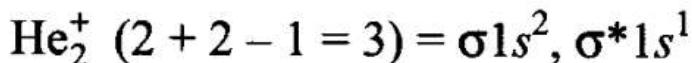
$$= \dfrac{4-4}{2} = 0$$

Here, bond order of He_2 is zero. Thus, it does not exist.

(b) He_2 $(2+2=4) = \sigma 1s^2, \sigma*1s^2$
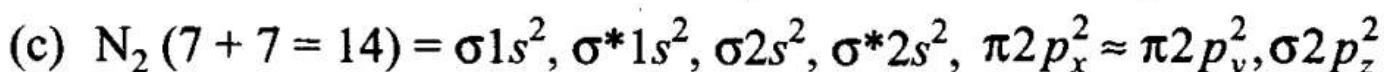
$$BO = \dfrac{2-2}{2} = 0$$

Here, bond order of He_2 is zero. Thus, it does not exist.

He_2^+ $(2+2-1=3) = \sigma 1s^2, \sigma*1s^1$

$$BO = \dfrac{2-1}{2} = 0.5$$
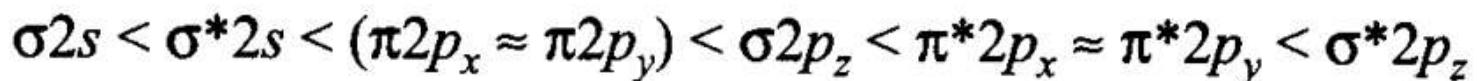
Since, the bond order is not zero, this molecule is expected to exist.

(c) N_2 $(7+7=14) = \sigma 1s^2, \sigma*1s^2, \sigma 2s^2, \sigma*2s^2, \pi 2p_x^2 \approx \pi 2p_y^2, \sigma 2p_z^2$

$$BO = \dfrac{10-4}{2} = 3$$

Thus, dinitrogen (N_2) molecule contains triple bond and no other molecule of second period have more than double bond. Hence, bond strength of N_2 is maximum amongst the homonuclear diatomic molecules belonging to the second period.
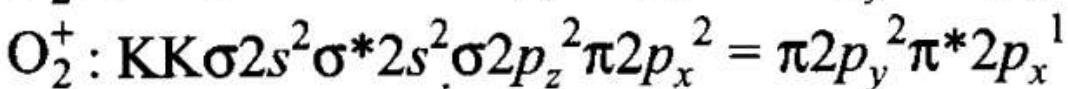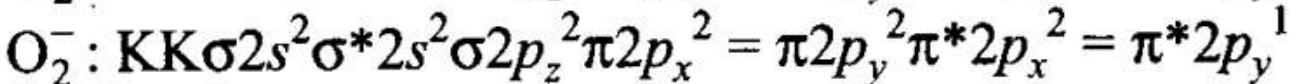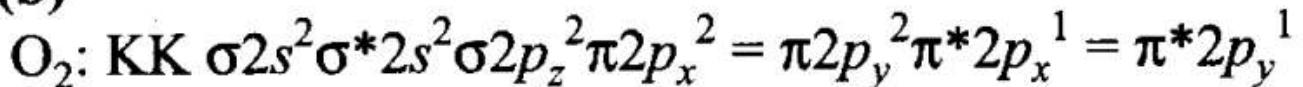
(d) It is incorrect. The correct order of energies of molecular orbitals in N_2 molecule is

$$\sigma 2s < \sigma*2s < (\pi 2p_x \approx \pi 2p_y) < \sigma 2p_z < \pi*2p_x \approx \pi*2p_y < \sigma*2p_z$$

Q20. Which of the following options represents the correct bond order?

(a) $O_2 > O_2 > O^+_2$

(b) $O^-_2 < O_2 < O_2$

(c) $O_2 > O_2 < O_2$

(d) $O^-_2 < O_2 > O^+_2$

Sol. (b)

$$O_2:\; KK\; \sigma 2s^2 \sigma^*2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^*2p_x^1 = \pi^*2p_y^1$$

$$O^-_2:\; KK\sigma 2s^2 \sigma^*2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^*2p_x^2 = \pi^*2p_y^1$$

$$O^+_2:\; KK\sigma 2s^2 \sigma^*2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^*2p_x^1$$

Bond order, O_2 $\qquad \dfrac{8-4}{2} = 2.0$

Bond order, O^-_2 $\qquad \dfrac{8-5}{2} = 1.5$

Bond order, O^+_2 $\qquad \dfrac{8-3}{2} = 2.5$

$\therefore \quad O^+_2 > O_2 > O^-_2 \quad$ or $\quad O^-_2 < O_2 < O^+_2$

Q21. The electronic configuration of the outermost shell of the most electronegative element is

(a) $2s^2 2p^5$

(b) $3s^2 3p^5$

(c) $4s^2 4p^5$

(d) $5s^2 5p^s$

Sol: (a) The electronic configuration represents

$2s^2 2p^5$ = fluorine = most electronegative element

$3s^2 3p^5$ = chlorine

$4s^2 4p^5$ = bromine

$5s^2 5p^5$ = iodine

Q22. Amongst the following elements, whose electronic configurations are given below, the one having the highest ionization enthalpy is

(a) $[Ne]3s^2 3p^1$

(b) $[Ne]3s^2 3p^3$

(c) $[Ne]3s^2 3p^2$

(d) $[Ar]3d^{10} 4s^2 4p^3$

Sol: (b) (b) and (d) have exactly half-filled p-orbitals but (b) is smaller in size than Hence, (b) has highest ionization enthalpy.

More than One Correct Answer Type

Q23. Which of the following have identical bond order?

(a) CN^- (b) NO^+ (c) O^-_2 (d) O_2^{2-}

Sol: (a, b) CN^- (number of electrons = 6 + 7 + 1 = 14)

NO^+ (number of electrons = 7 + 8 − 1 = 14)

O^-_2 (number of electrons = 8 + 8 + 1 = 17)

O_2^{2-} (number of electrons = 8 + 8 + 2 = 18)

Thus, CN and NO^+ because of the presence of same number of electrons, have same bond order.

Q24. Which of the following attain the linear structure?

(a) $BeCl_2$

(b) NCO^+

(c) NO_2

(d) CS_2

Sol: (a, d) $BeCl_2$ (Cl − Be − Cl) and CS_2 (S = C = S) both are linear, NCO^+ is non-linear. However, remember that NCO(N = C = O) is linear because it is isoelectronic with CO_2.

NO_2 is angular with bond angled 132° and each O − N bond length of 1.20A° (intermediate between single and double bond).

Q25. CO is isoelectronic with

(a) NO^+

(b) N_2

(c) $SnCl_2$

(d) NO_2

Sol: (a, b) Number of electrons in CO =14

Number of electrons in NO^+ =14

Number of electrons in N_2 = 14
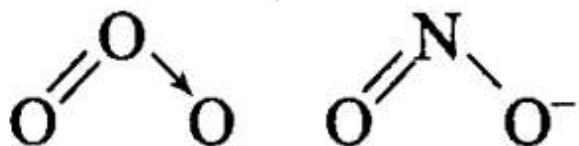
Number of electrons in $SnCl_2$ = 84

Number of electrons in NO^-_2 = 24

Q26. Which of the following species have the same shape?

(a) CO_2

(b) CCl_4

(c) O_3

(d) NO^-_2

Sol: (c, d) $CO_2 \rightarrow$ Linear, $CCl_4 \rightarrow$ Tetrahedral, $O_3 \rightarrow$ Angular (V-shaped), $NO^-_2 \rightarrow$ Angular (V-shaped)

Q27. Which of the following statements are correct about CO_3^{2-} ?

(a) The hybridization of central atom is sp^3.

(b) Its resonance structure has one C – O single bond and two C = O double bonds.

(c) The average formal charge on each oxygen atom is 0.67 units.

(d) All C – O bond lengths are equal.

Sol. (c, d)
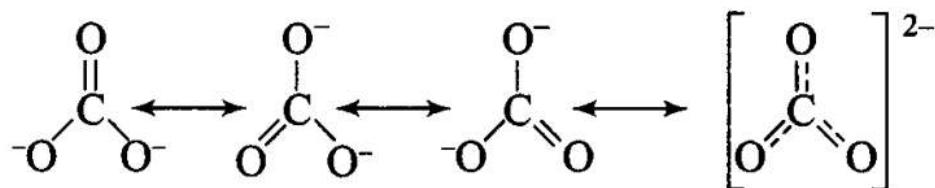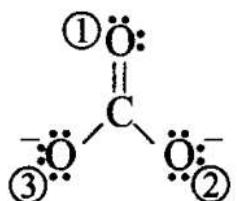
Formal charge on O atom (1): $6 - 4 - \dfrac{1}{2} \times 4 = 0$

Formal charge on O atom (2): $6 - 6 - \dfrac{1}{2} \times 2 = -1$

Formal charge on O atom (3): $6 - 6 - \dfrac{1}{2} \times 2 = -1$

Average formal charge on each oxygen atom: $\dfrac{-1 + (-1) + 0}{3} = -0.67$

All C – O bonds are equal due to resonance.

Q28. Diamagnetic species are those which contain no unpaired electrons. Which among the following are diamagnetic?

(a) N_2

(b) N_2^{2-}

(c) O_2

(d) O_2^{2-}

Sol. (a, d)

(a) Electronic configuration of $N_2 = \sigma 1s^2 \; \sigma^*1s^2 \; \sigma 2s^2 \; \sigma^*2s^2 \; \pi 2p_x^2 \approx \pi 2p_y^2 \; \sigma 2p_z^2$. It has no unpaired electron. This indicates it is a diamagnetic species.

(b) Electronic configuration of N_2^{2-} ion $= \sigma 1s^2 \; \sigma^*1s^2 \; \sigma 2s^2 \; \sigma^*2s^2$
$\pi 2p_x^2 \approx \pi 2p_y^2 \; \sigma 2p_z^2 \; \pi 2p_x^1 \approx \pi^*2p_y^1$.
It has two unpaired electrons, so it is paramagnetic in nature.

(c) Electronic configuration of $O_2 = \dot{\sigma}1s^2 \; \sigma^*1s^2 \; \sigma 2s^2 \; \sigma^*2s^2$

$\sigma 2p_z^2 \; \pi 2p_x^2 \approx \pi 2p_y^2 \; \pi^*2p_x^1 \approx \pi^*2p_y^1$.

The presence of two unpaired electrons shows its paramagnetic nature.

(d) Electronic configuration of O_2^{2-} ion $= \sigma 1s^2 \; \sigma^*1s^2 \; \sigma 2s^2 \; \sigma^*2s^2$

$\sigma 2p_z^2 \; \pi 2p_x^2 \approx \sigma 2p_y^2 \; \pi^*2p_x^2 \approx \pi^*2p_y^2$.

It contains no unpaired electron, therefore, it is diamagnetic in nature.

Q29. Species having same bond order are

(a) N_2

(b) N^-_2

(c) F^+_2

(d) O^-_2

Sol. (c, d) The total number of electrons in F_2^+ and O_2^- are same, i.e. 17 electrons.

$$F_2^+ (17): \text{B.O.} = \frac{10-7}{2} = 1.5$$

$$O_2^- (17): \text{B.O.} = \frac{10-7}{2} = 1.5$$

Q30. Which of the following statements are not correct?

(a) NaCl being an ionic compound is a good conductor of electricity in the solid state.

(b) In canonical structures there is a difference in the arrangement of atoms.

(c) Hybrid orbitals form stronger bonds than pure orbitals.

**(d) VSEPR theory can explain the square planar geometry of XeF$_4$.**

Sol: (a, b)
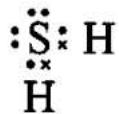
(a) Ionic compounds are good conductors only in molten state or aqueous solution since ions are not furnished in solid state.

(b) In canonical structures there is a difference in arrangement of electrons.
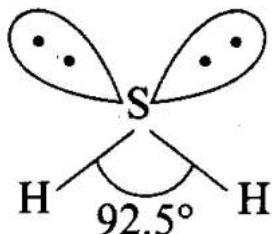
Short Answer Type Questions

Q31. Explain the non linear **shape of H$_2$S and** non planar **shape of PCl$_3$ using valence shell electron pair repulsion theory.**
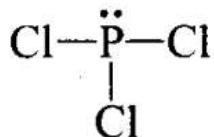
Sol: The Lewis structure of H$_2$S is:

$$:\overset{\cdot\cdot}{\underset{\bullet\times}{S}}\times H$$
$$H$$
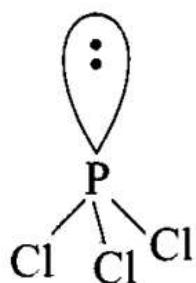
S-atom is surrounded by four electron pairs (two bonded and two lone pairs). These four electron pairs adopt tetrahedral arrangement. The presence of two lone pairs brings distortion in the molecule on account of repulsion with bonded pairs of electrons. Thus, the shape of H_2S molecule is V-shaped and not linear.


The Lewis structure of PCl_3 is:

$$Cl-\overset{\cdot\cdot}{P}-Cl$$
$$|$$
$$Cl$$

P-atom is surrounded by four electron pairs (3 bonded and one lone pair). These four pairs adopt a tetrahedral geometry. Due to the presence of lone pair, PCl_3 has a distorted tetrahedral geometry. Thus, it is pyramidal in shape and not non-planar shape.

Q32. Using molecular orbital theory, compare the bond energy and magnetic character of O_2^+ and O_2^-

Sol. O_2^+ (15): $\sigma 1s^2 \; \sigma^*1s^2 \; \sigma 2s^2 \; \sigma^*2s^2 \; \sigma 2p_z^2 \; \pi 2p_x^2 = \pi 2p_y^2 \; \pi^*2p_x^1$

$$\text{B.O.} = \frac{1}{2}(10 - 5) = 2.5$$

O_2^- (17): $\sigma 1s^2 \; \sigma^*1s^2 \; \sigma 2s^2 \; \sigma^*2s^2 \; \sigma 2p_z^2 \; \pi 2p_x^2 = \pi 2p_y^2 \; \pi^*2p_x^2 = \pi^*2p_y^1$

$$\text{B.O.} = \frac{1}{2}(10 - 7) = 1.5$$

Bond energy of $O_2^+ > O_2^-$
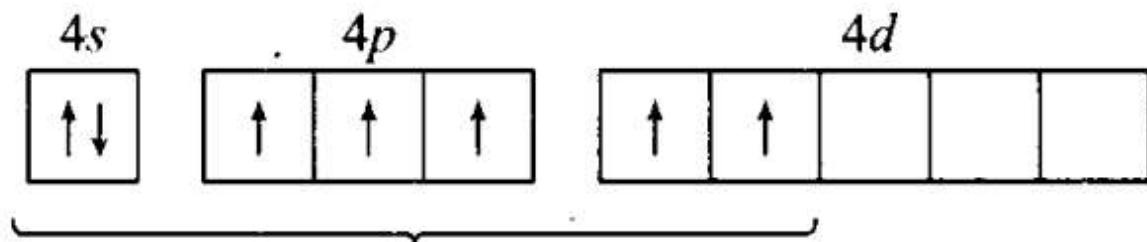
Both are paramagnetic due to presence of unpaired electrons.

Q33. Explain the shape of BrF$_5$.

Sol: Br-atom has configuration:
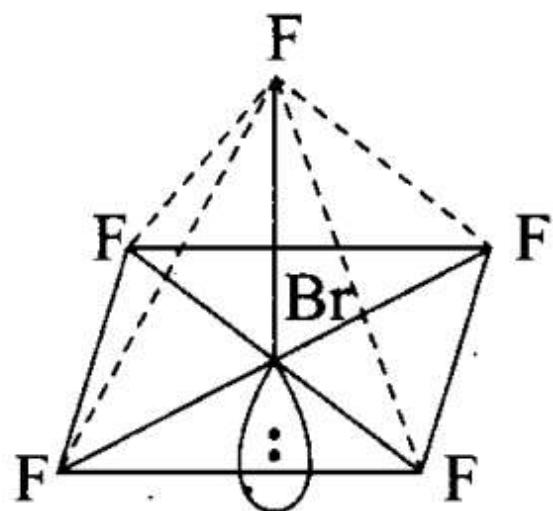
$1s^2, \; 2s^2 2p^6, \; 3s^2 3p^6 3d^{10}, \; 4s^2 4p^5$

To get pentavalency, two of the p-orbitals are unpaired and electrons are shifted to 4d-orbitals.
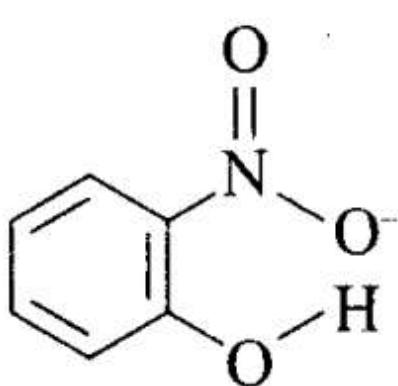

In this excited state, sp^{3}d^2-hybridisation occurs giving octahedral structure. Five positions are occupied by F atoms forming sigma bonds with hybrid bonds and one position occupied by lone pair, i.e., the molecule as
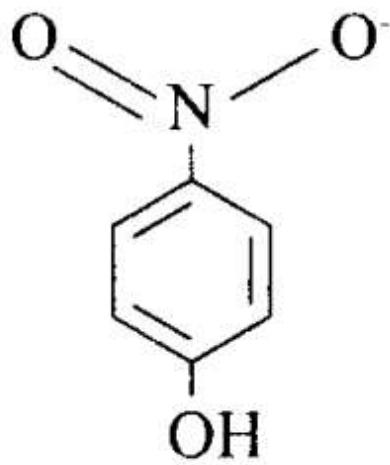
a square pyramidal shape.

Q34. Structures of molecules of two compounds are given below:
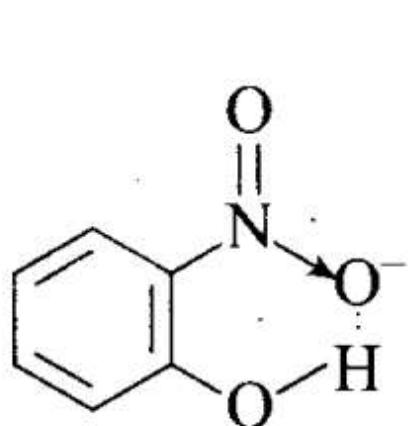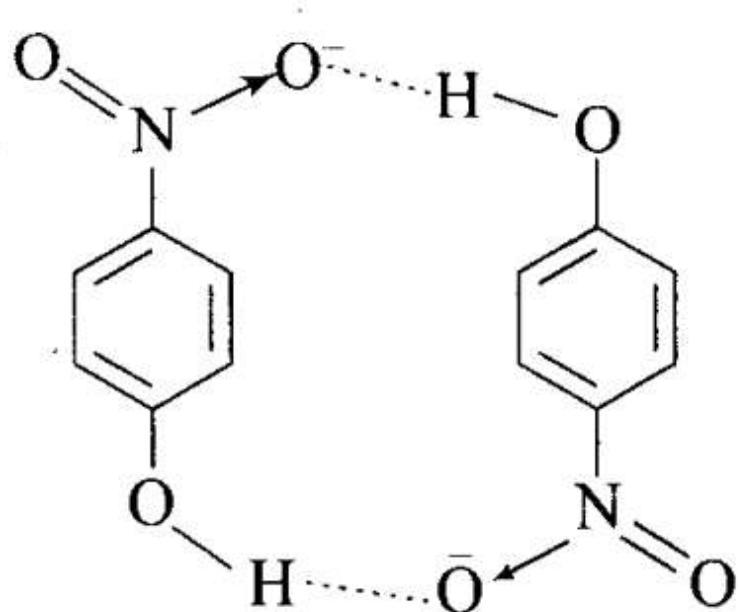

(i)   (ii)

(a) Which of the two compounds will have intermolccular hydrogen bonding and which compound is expected to show intramolecular hydrogen bonding?
(b) The melting point of a compound depends on. among other things, the extent of hydrogen bonding. On this basis explain which of the above two compounds will show higher melting point.
(c) Solubility of compounds in water depends on power to form hydrogen bonds with water. Which of the above compounds will form hydrogen bond with water easily and be more soluble in it?
Sol:(a) Compound I will form intramolecular hydrogen bond because N02 and OH groups are close together

whereas it is not so in compound II. Compound II will have intermolecular hydrogen bonding. Bonding in both the cases is shown below:
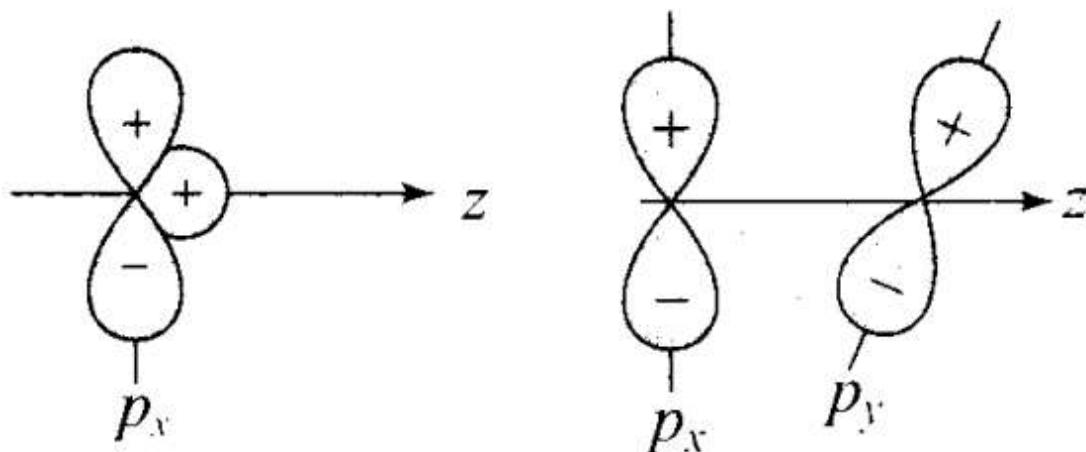
(i)   (ii)

(b) As a large number of molecules can be linked together through intermolecular hydrogen bonding, compound II will show higher melting point.

(c) Due to intramolecular hydrogen bonding, compound I will not be able to form hydrogen bonds with H_2O molecules. Hence, it will be less soluble in water. However, compound II can form hydrogen bonds with H_2O molecules easily and hence it will be more soluble in water.

Q35. Why does type of overlap given in the following figure not result in the bond formation?
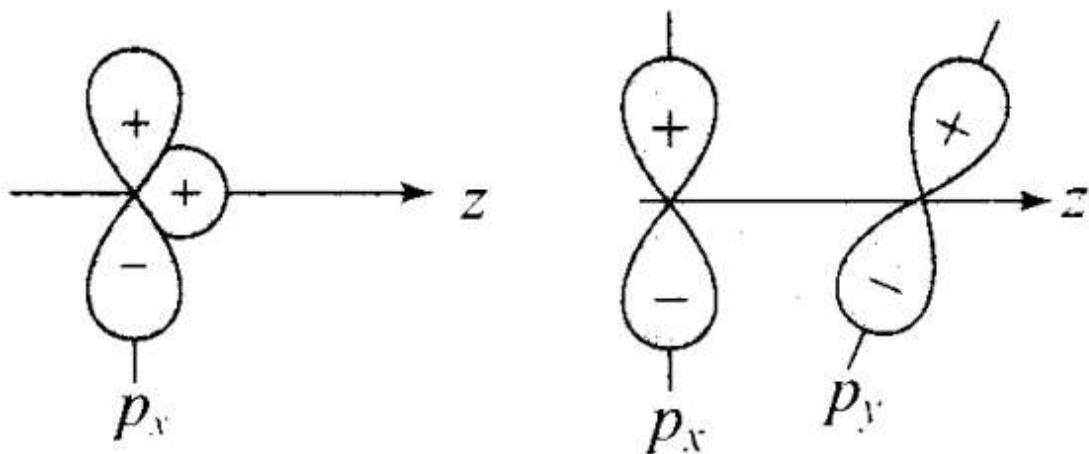
Sol: In first figure, the ++ overlap is equal to +- overlap and therefore, these cancel out and net overlap is zero.

In second figure, no overlap is possible because the two orbitals are perpendicular to each other.
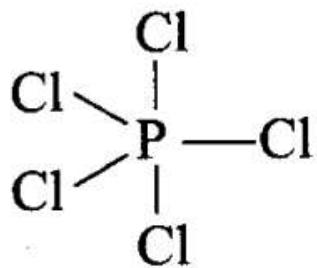
**Q36. Explain why PC1$_5$ is trigonal bipyramidal whereas IF$_5$ is square pyramidal.**

Sol: PC1$_5$ – The ground state and the excited state outer electronic configurations of phosphorus (Z = 15) are represented below
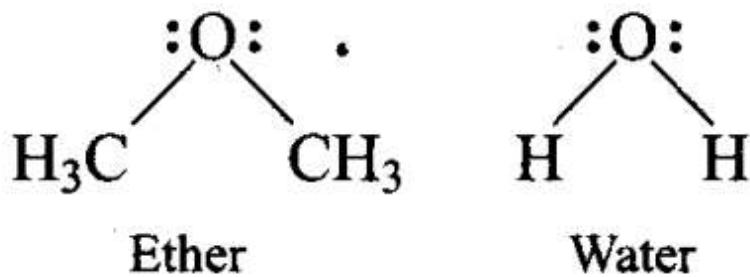
P(ground state) 1↿⇂ | 1 | 1 | 1 | | | | | |
  3s   3p   3d

P(excited state) | 1 | | 1 | 1 | 1 | | 1 | | | | |
                     sp^3d

37. In both water and dimethyl ether $(CH_3 - \overset{..}{\underset{..}{O}} - CH_3)$, oxygen atom is central atom; and has the same hybridization, yet they have different bond angles. Which one has greater bond angle? Give reason.
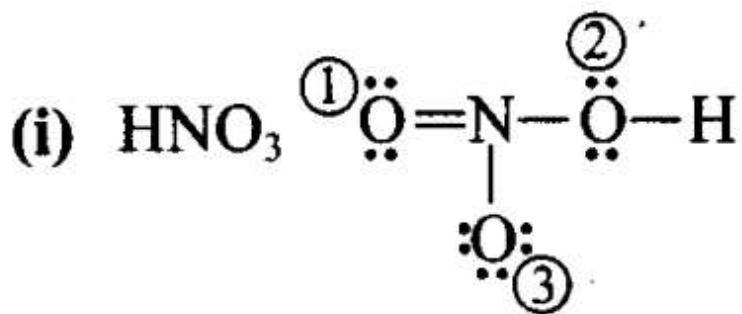
Sol: Dimethyl ether has larger bond angle than water. This is because there is more repulsion between bond pairs of CH_3 groups attached in ether than between bond pairs of hydrogen atoms attached to oxygen in water. The carbon of CH_3 group in ether is attached to three hydrogen atoms through c bonds and electron pairs of these bonds add to the electron charge density on carbon atom. Hence, repulsion between two - CH_3 groups will be more than that between two H atoms.

Ether   Water

Q38. Write Lewis structure of the following compounds and show formal charge on each atom. HNO_3, NO_2, H_2SO_4

Sol: Formal charge on an atom in a Lewis structure

= [total number of valence electrons in free atom] − [total number of non-bonding (lone pairs) electrons]

−1/2 [total number of bonding or shared electrons]

(i) HNO_3

$$\overset{\textcircled{1}}{\ddot{O}}=N-\overset{\textcircled{2}}{\underset{|}{\ddot{O}}}-H$$
$$:\overset{\cdot\cdot}{\underset{\textcircled{3}}{O}}:$$
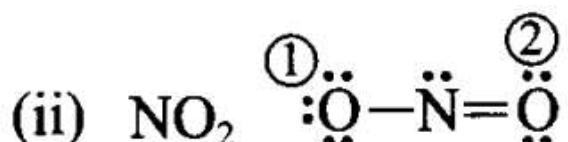
F.C. on O (1) $\rightarrow 6 - 4 - \dfrac{1}{2}(4) = 0$

F.C. on O (2) $\rightarrow 6 - 4 - \dfrac{1}{2}(4) = 0$

F.C. on O (3) $\rightarrow 6 - 6 - \dfrac{1}{2}(2) = -1$

F.C. on N $\rightarrow 5 - 0 - \dfrac{1}{2}(8) = +1$

F.C. on H $\rightarrow 1 - 0 - \dfrac{1}{2}(2) = 0$

(ii) NO_2

$$\overset{\textcircled{1}}{:\ddot{O}}-\ddot{N}=\overset{\textcircled{2}}{\ddot{O}}$$
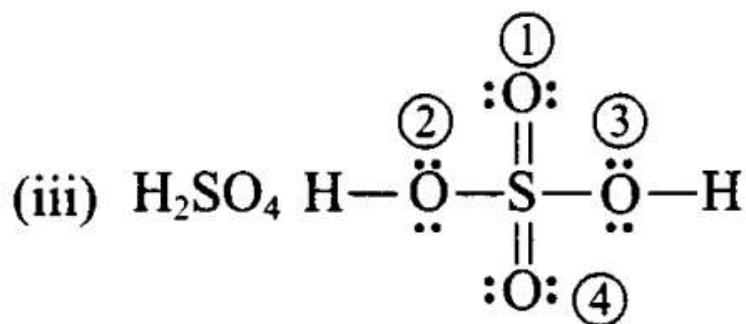
F.C. on O (1) $\rightarrow 6 - 6 - \dfrac{1}{2}(2) = -1$

F.C. on O (2) $\rightarrow 6 - 4 - \dfrac{1}{2}(4) = 0$

F.C. on N $\rightarrow 5 - 2 - \dfrac{1}{2}(6) = 0$

(iii) H_2SO_4

$$\text{①} \quad :\overset{\text{①}}{\underset{}{O}}: $$

$$\text{(iii)} \quad H_2SO_4 \quad H-\overset{\text{②}}{\underset{\cdot\cdot}{O}}-\overset{}{\underset{}{S}}-\overset{\text{③}}{\underset{\cdot\cdot}{O}}-H$$

$$:\overset{}{\underset{\text{④}}{O}}:$$
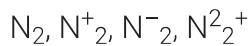
F.C. on O (1) and (4) $\rightarrow 6 - 4 - \dfrac{1}{2}(4) = 0$

F.C. on O (2) and (3) $\rightarrow 6 - 4 - \dfrac{1}{2}(4) = 0$
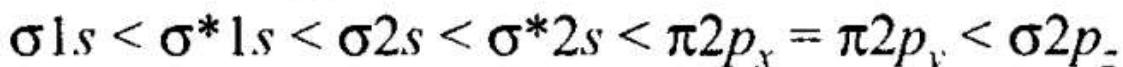
F.C. on H $\longrightarrow 1 - 0 - \dfrac{1}{2}(2) = 0$

F.C. on S $\longrightarrow 6 - 0 - \dfrac{1}{2}(12) = 0$
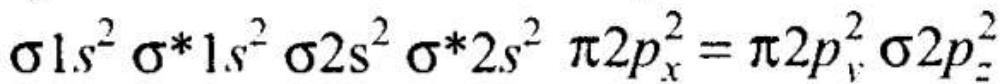
**Q39.** The energy of $\sigma 2p_z$ molecular orbital is greater than $2p_x$ and $2p_y$ molecular orbitals in nitrogen molecule. Write the complete sequence of energy levels in the increasing order of energy in the molecule. Compare the relative stability and the magnetic behaviour of the following species:
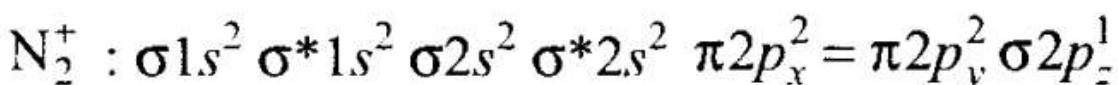
$N_2, N^+_2, N^-_2, N^{2+}_2$

Sol. Sequence of energy levels

$$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \pi 2p_x = \pi 2p_y < \sigma 2p_z$$
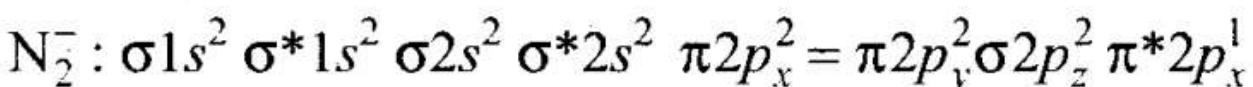
For N_2 molecule the M.O. configuration is:

$$\sigma 1s^2 \; \sigma^* 1s^2 \; \sigma 2s^2 \; \sigma^* 2s^2 \; \pi 2p_x^2 = \pi 2p_y^2 \; \sigma 2p_z^2$$

$$\text{B.O.} = \frac{1}{2}(10 - 4) = 3, \text{ diamagnetic}$$

$$N_2^+ : \sigma 1s^2 \; \sigma^* 1s^2 \; \sigma 2s^2 \; \sigma^* 2s^2 \; \pi 2p_x^2 = \pi 2p_y^2 \; \sigma 2p_z^1$$

$$\text{B.O.} = \frac{1}{2}(9 - 4) = 2.5, \text{ paramagnetic}$$

$$N_2^- : \sigma 1s^2 \; \sigma^* 1s^2 \; \sigma 2s^2 \; \sigma^* 2s^2 \; \pi 2p_x^2 = \pi 2p_y^2 \sigma 2p_z^2 \; \pi^* 2p_x^1$$
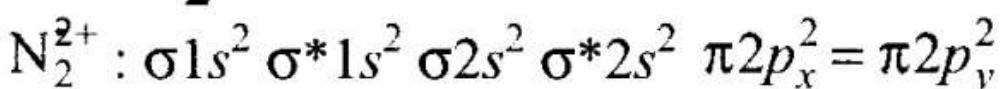
$$\text{B.O.} = \frac{1}{2}(10 - 5) = 2.5, \text{ paramagnetic}$$

$$N_2^{2+} : \sigma 1s^2 \; \sigma^* 1s^2 \; \sigma 2s^2 \; \sigma^* 2s^2 \; \pi 2p_x^2 = \pi 2p_y^2$$

$$\text{B.O.} = \frac{1}{2}(8 - 4) = 2, \text{ diamagnetic}$$

Stability order : $N_2 > N_2^- > N_2^+ > N_2^{2+}$

(N_2^- has more bonding electrons as compared to N_2^+).

Q40. What is the effect of the following processes on the bond order in N-, and O_2?
(i) $N_2 \rightarrow N^+{}_2 + e^-$ (ii) $O_2 \rightarrow O^+{}_2 + e^-$

Sol.
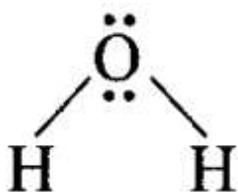
| Species	Total electrons	Configuration	Bond order
N_2	14	KK $\sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2\ \sigma 2p_z^2$	$\dfrac{8-2}{2} = 3$
N_2^+	13	KK $\sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2\ \sigma 2p_z^1$	$\dfrac{7-2}{2} = 2.5$
Bond order decreases from 3 to 2.5			
O_2	16	KK $\sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2\ \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^1 = \pi^* 2p_y^1$	$\dfrac{8-4}{2} = 2.0$
O_2^+	15	KK $\sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2\ \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^1$	$\dfrac{8-3}{2} = 2.5$
Bond order increases from 2 to 2.5			

Q41. Give reasons for the following: '

(a) Covalent bonds are directional bonds while ionic bonds are non- directional.

(b) Water molecule has bent structure whereas carbon dioxide molecule is linear.

(c) Ethyne molecule is linear.
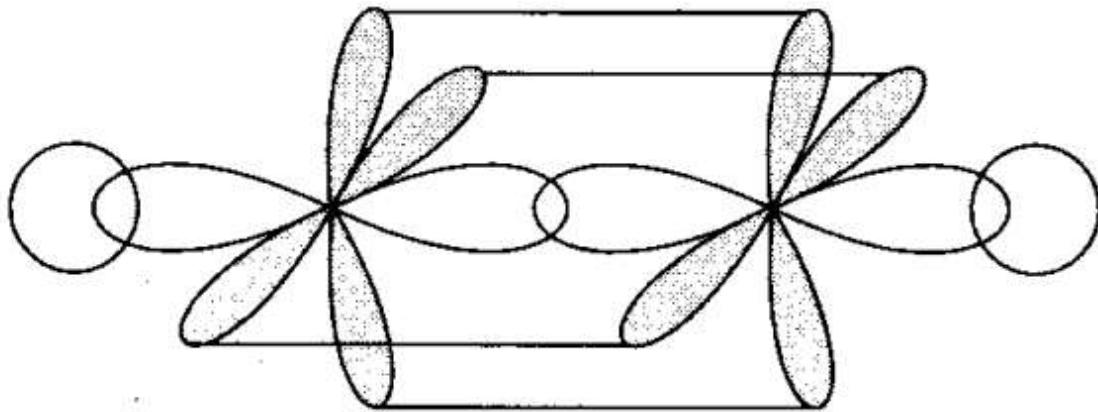
Sol: (i) Since the covalent bond depends on the overlapping of orbitals between different orbitals, the geometry of the molecule is different. The orientation of overlap is different. The orientation of overlap is the factor responsible for their directional nature.

(ii) Due to presence of two lone pairs of electrons on oxygen atom in HiO the repulsion between lp-lp is more. CO_2 undergoes sp hybridization resulting in linear shape (O = C = O) while H_2O undergoes .sp^3 hybridisation resulting in distorted tetrahedral or bent structure.
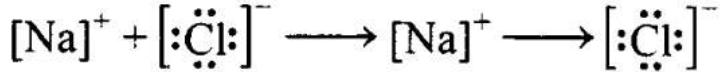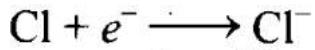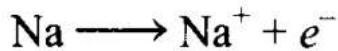

(iii) In ethyne molecule carbon undergoes sp hybridization with two unhybridised orbitals. One sp hybrid orbital of one carbon atom overlaps axially with sp hybrid orbital of the other carbon atom to form C – C

sigma bond while the other hybridized orbital of each carbon atom overlaps axially with S orbitals of hydrogen atoms forming σ bonds. Unhybridised orbitals form π bonds
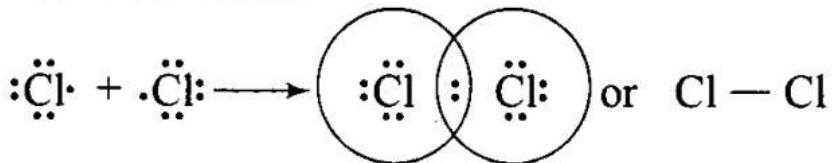

Q42. What is an ionic bond? With two suitable examples explain the difference between an ionic and covalent bond?

Sol: An ionic bond is formed as a result of the electrostatic attraction between the positive and negative ions formed by transfer of electrons from one atom to another.

$$Na \longrightarrow Na^+ + e^-$$
$$Cl + e^- \longrightarrow Cl^-$$
$$[Na]^+ + \left[:\ddot{\underset{..}{C}}l:\right]^- \longrightarrow [Na]^+ \longrightarrow \left[:\ddot{\underset{..}{C}}l:\right]^-$$
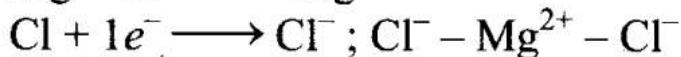
Covalent bond is formed by sharing of electrons between atoms. Examples of covalent bonds:

Multiple bonds are formed during covalent bonding while in ionic bond a number of bonds can be formed between different atoms.

$$Mg - 2e^- \longrightarrow Mg^{2+}$$
$$Cl + 1e^- \longrightarrow Cl^- ; Cl^- - Mg^{2+} - Cl^-$$

Q43. Arrange the following bonds 'in order of increasing ionic character giving reason.
N-H, F-H, C-H and O-H

Sol. Electronegativity difference
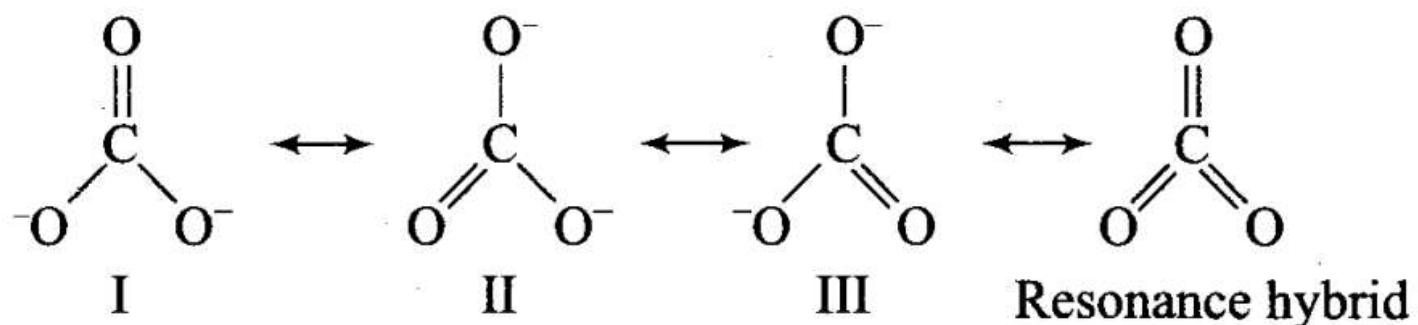
$$N\text{—}H \qquad F\text{—}H \qquad C\text{—}H \qquad O\text{—}H$$
$$3.0 - 2.1 \qquad 4.0 - 2.1 \qquad 2.5 - 2.1 \qquad 3.5 - 2.1$$
$$= 0.9 \qquad = 1.19 \qquad = 0.4 \qquad = 1.4$$

Increasing order of electronegativity difference: C-H<N-H<0-H<F-H

Greater is the difference in electronegativity between two bonded atoms, greater is the ionic character.
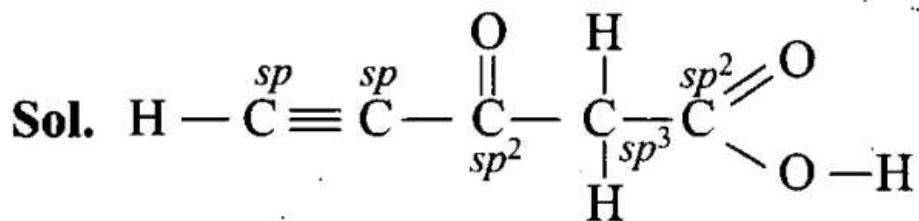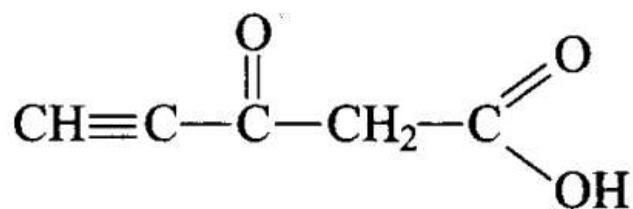
Q44. Explain why CO_3^{2-} ion cannot be represented by a single Lewis structure. How can it be best represented?

Sol: A single Lewis structure of CO_3^{2-} ion cannot explain all the properties of this ion. It can be represented as a resonance hybrid of the following structures:

I   II   III   Resonance hybrid

If it were represented only by one structure, there should be two types of bonds, i.e., C = O double bond and C – O single bonds but actually all bonds are found to be identical with same bond length and same bond strength.

Q45. Predict the hybridization of each carbon in the molecule of organic compound given below. Also indicate the total number of sigma and pi bonds in this molecule.

$$\underset{\substack{}}{CH\equiv C-\overset{\overset{\textstyle O}{\|}}{C}-CH_2-\overset{\textstyle O}{\underset{\textstyle OH}{C}}}$$

Sol. $H-\underset{}{\overset{sp}{C}}\equiv\overset{sp}{C}-\underset{sp^2}{\overset{\overset{\textstyle O}{\|}}{C}}-\underset{H}{\overset{H}{\underset{|}{\overset{|}{C}}}}\underset{sp^3}{}-\underset{O-H}{\overset{sp^2}{\overset{\textstyle O}{C}}}$
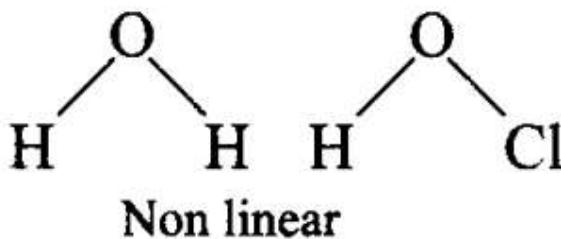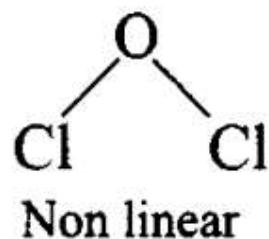
Number of σ bonds = 11
Number of π bonds = 4

Q46. Group the following in linear and non-linear molecules: H_2O, HOC1, $BeCl_2$ Cl_2O

Sol.
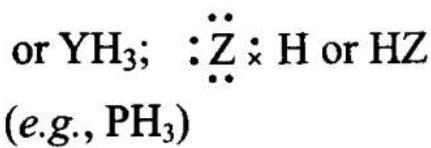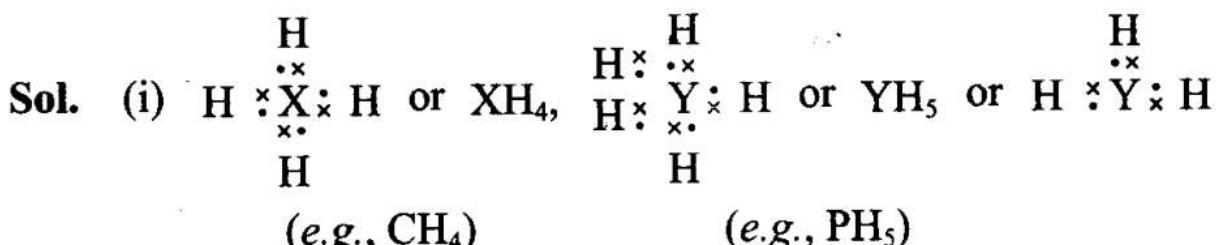
Q47. Elements X, Y and Z have 4, 5 and 7 valence electrons respectively, (i) Write the molecular formula of the compounds formed by these elements individually with hydrogen, (ii) Which of these compounds will have the highest dipole moment?

Sol. (i) H ⁚X⁚ H or XH_4, H⁚ Y⁚ H or YH_5 or H ⁚Y⁚ H
(with H above and below X and Y)

(e.g., CH_4)   (e.g., PH_5)

or YH_3; ⁚Z⁚ H or HZ

(e.g., PH_3)
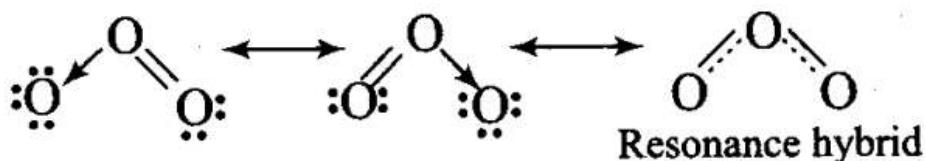
(ii) Z will be most electronegative element and hence HZ will have highest dipole moment.
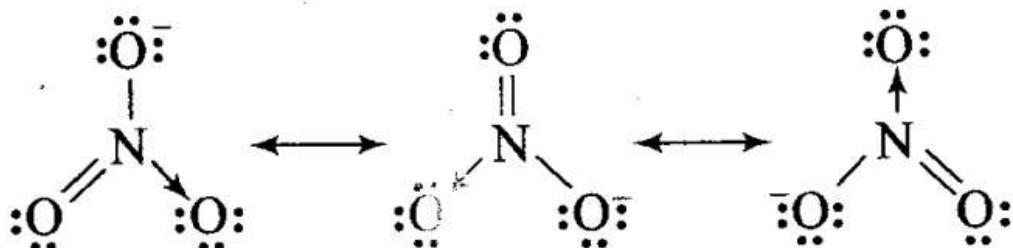
Q48. Draw the resonating structure of (i) Ozone molecule (ii) Nitrate ion

Sol. Resonating structures:

    (i) Ozone molecule

Resonance hybrid

    (ii) Nitrate ion

Q49. Predict the shapes of the following molecules on the basis of hybridization. BCl_3, ch_4, co_2, nh_3
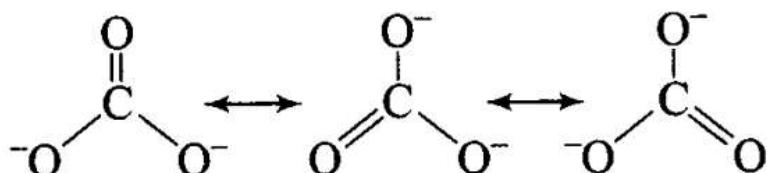
Sol: BCl_3 – sp^2 hybridisation – Trigonal planar

CH_4 – sp^3 hybridisation – Tetrahedral .
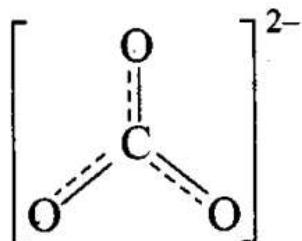
NH_3 – sp^3 hybridisation – Distorted tetrahedral or Pyramidal

Q50. All the C – O bonds in carbonate ion (CO^{2-}_3) are equal in length. Explain.

Sol: Carbonate ion is represented by resonating structures as given below:
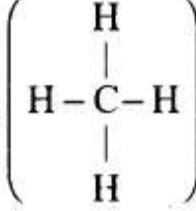
Actual structure will be the mixture of all resonating structures, also called resonance hybrid structure.

Q51. What is meant by the term average bond enthalpy? Why there is difference in bond enthalpy of O – H bond in ethanol (C_2H_5OH) and water?

Sol. All the similar bonds in a molecule do not have the same bond enthalpies,

e.g., in CH_4 $\left(\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}\right)$ the four C – H bonds do not have the same bond

enthalpies because after breaking of the bonds one by one, the electronic environment around carbon changes. Hence, the actual bond enthalpy of C – H bond is taken as the average value.

O – H bond in ethanol $\left(C_2H_5\!-\!O_{\diagdown H}\right)$ and that in water $\left(H\!-\!O_{\diagdown H}\right)$ do not have similar electronic

Environment around oxygen atom. Hence, their O – H bond enthalpies are different.

Matching Column Type Questions

Q52. Match the species in Column I with the type of hybrid orbitals in Column II.

Column I	Column II
(i) SF_4	(a) sp^3cf
(ii) if_5	(b) d^2sp^3
(iii) $NO_2{}^+$	(c) $sp^3 d$
(iv) NH_4	(d) sp3
	(e) sp

Sol. (i) → c; (ii) → a; (iii) → e; (iv) → d

(i) SF_4 = number of bp (4) + number of lp (1)
= sp^3d hybridisation

(ii) IF_5 = number of bp (5) + number of lp (1)
= sp^3d^2 hybridisation

(iii) NO_2^+ = number of bp (2) + number of lp (0)
= sp hybridisation

(iv) NH_4^+ = number of bp (4) + number of lp (0)
= sp^3 hybridisation

Q53. Match the species in Column I with the geometry/shape in Column II.

Column I	Column II
(i) H_3O^+	(a) Linear
(ii) HC = CH	(b) Angular
(iii) $ClO^-{}_2$	(c) Tetrahedral
(iv) $NH^+{}_4$	(d) Trigonal bipyramidal
—	(e) Pyramidal

Sol. (i) → e; (ii) → a; (iii) → b; (iv) → c

H_3O^+ – pyramidal

$CH \equiv CH$ – linear

ClO_2^- -- angular

NH_4^+ – tetrahedral

Q54. Match the species in Column I with the bond order in Column II.

Column I	. Column II
(i) NO	(a) 1.5
(ii) CO	(b) 2.0
(iii) O^-_2	(c) 2.5
(iv) O_2	(d) 3.0

Sol. (i) → c; (ii) → d; (iii) → a; (iv) → b

$$NO \rightarrow B.O. = \frac{1}{2}(10 - 5) = 2.5$$

$$CO \longrightarrow B.O. = \frac{1}{2}(10 - 4) = 3$$

$$O^-_2 \cdot \longrightarrow B.O. = \frac{1}{2}(10 - 7) = 1.5$$

$$O_2 \longrightarrow B.O. = \frac{1}{2}(10 - 6) = 2$$

Q55. Match the items given in Column I with examples given in Column II.

| Column I	Column II
(i) Hydrogen bond	(a) C
(ii) Resonance	(b) LiF
(iii) Ionic solid	(c) H_2
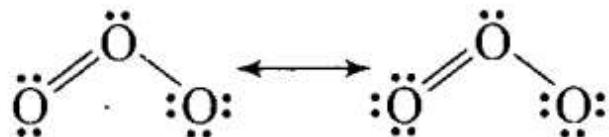
(iv) Covalent solid	(d) HF
	(e) O_3

Sol. (i) → d; (ii) → e; (iii) → b; (iv) → a

   (i) H − F : hydrogen bond H − F ………. H − F

  (ii) O_3 : Resonance

 (iii) LiF: Ionic solid
 (iv) C: Covalent solid

Q56. Match the shape of molecules in Column I with the type of hybridization in Column II.

Column I	Column II
(i) Tetrahedral	(a) sp^2
(ii) Trigonal	(b) sp
(iii) Linear	(c) sp^3

Sol: (i) →c; (ii) → a; (iii) —> b

sp^3 hybridisation − Tetrahedralshape
sp^2 hybridisation − Trigonal shape
sp hybridization − Linear shape

Assertion and Reason Type Questions

In the following questions a statement of Assertion (A) followed by a statement of Reason (R) is given.
Choose the correct option out of the choices given below each question.

Q57. Assertion (A): Sodium chloride formed by the action of chlorine gas on sodium metal is a stable compound.

Reason (R): This is because sodium and chloride ions acquire octet in sodium chloride formation.

(a) A and R both are correct, and R is the correct explanation of A.

(b) A and R both are correct, but R is not the correct explanation of A.

(c) A is true but R is false.

(d) A and R both are false.
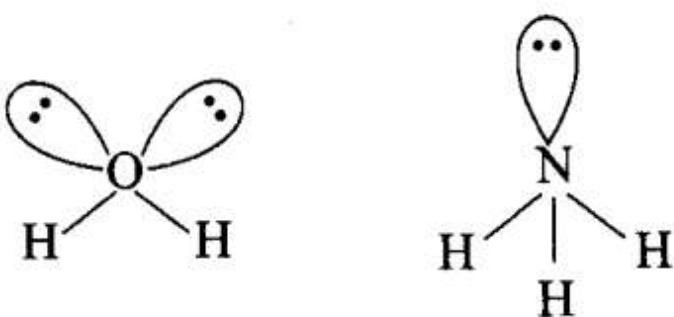
Sol: (a) Sodium chloride (Na^+CL^-) is stable ionic compound because both Na^+ and CL^- ions have complete octet in outermost shell.

Q58. Assertion (A): Though the central atom of both NH_3 and H_2O molecules are sp^3 hybridised, yet $H - N - H$ bond angle is greater than that of $H - O - H$.

Reason (R): This is because nitrogen atom has one lone pair and oxygen atom has two lone pairs.

(a) A and R both are correct, and R is the correct explanation of A.

(b) A and R both are correct, but R is not the correct explanation of A.

(c) A is true but R is false.

(d) A and R both are false.

Sol:(a)

H_2O has two lone pairs while NH_3 has single lone pair, hence, H_2O involves greater lone pair-bond pair repulsion.

Q59. Assertion (A): Among the two $O - H$ bonds in H_2O molecule, the energy required to break the first $O - H$ bond and other $O - H$ bond is the same.

Reason (R): This is because the electronic environment around oxygen is the same even after breakage of one $O - H$ bond.

(a) A and R both are correct, and R is the correct explanation of A.

(b) A and R both are correct, but R is not the correct explanation of A.

(c) A is true but R is false.

(d) A and R both are false.

Sol: (d) Bond energy of two (-O − H) bonds in H_2O will be different.

Long Answer Type Questions

Q60. (i) Discuss the significance/applications of dipole moment.

(ii) Represent diagrammatically the bond moments and the resultant dipole moment in CO_2, NF_3 and $CHCl_3$.

Sol: (i) Dipole moment plays very important role in understanding the nature of chemical bond. A few applications are given below:

(a) Distinction between, polar and non-polar molecules. The measurement of dipole moment can help us to distinguish between polar and non-polar molecules. Non-polar molecules have zero dipole moment while polar molecules have some value of dipole moment.
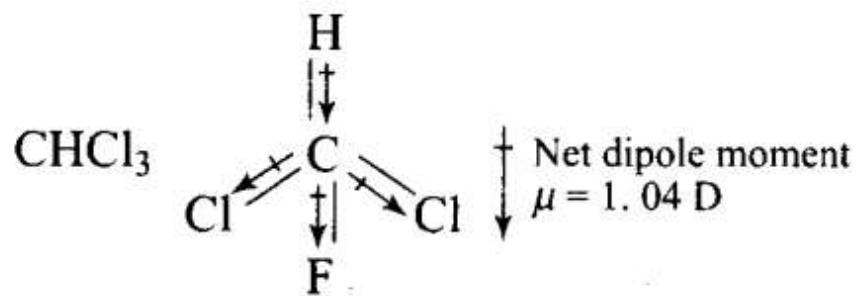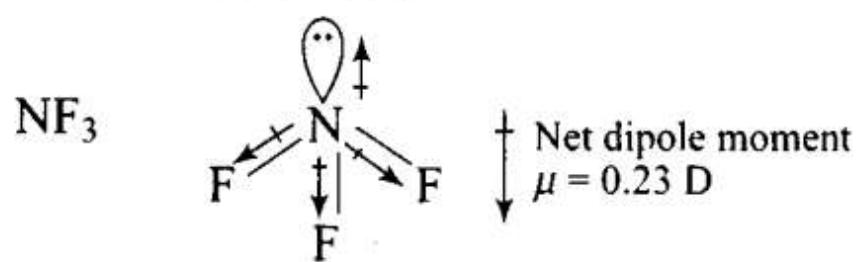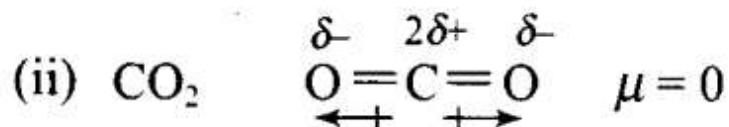
(b) Degree of polarity in a molecule. Dipole moment measurement also gives an idea about the degree of polarity specially in a diatomic molecule. The greater the dipole moment, the greater is the polarity in such a molecule.

(c) Shape of molecules. In case of molecules containing more than two atoms, the dipole moment not only depends upon the individual dipole moments of the bonds but also on the arrangement of bonds. .Thus, dipole moment is used to find the shapes of molecules.

(d) Ionic character in a molecule. Knowing the electronegativities of atoms involved in a molecule, it is possible to predict the nature of chemical bond formed. If the difference in electronegativities of two atoms is large, the bond will be highly polar. As an extreme case, when the electron is completely transferred from one atom to another, an ionic bond is formed. Therefore, the ionic bond is regarded as an extreme case of covalent bond. The greater the difference in electronegativities of the bonded atoms, the higher is the ionic character
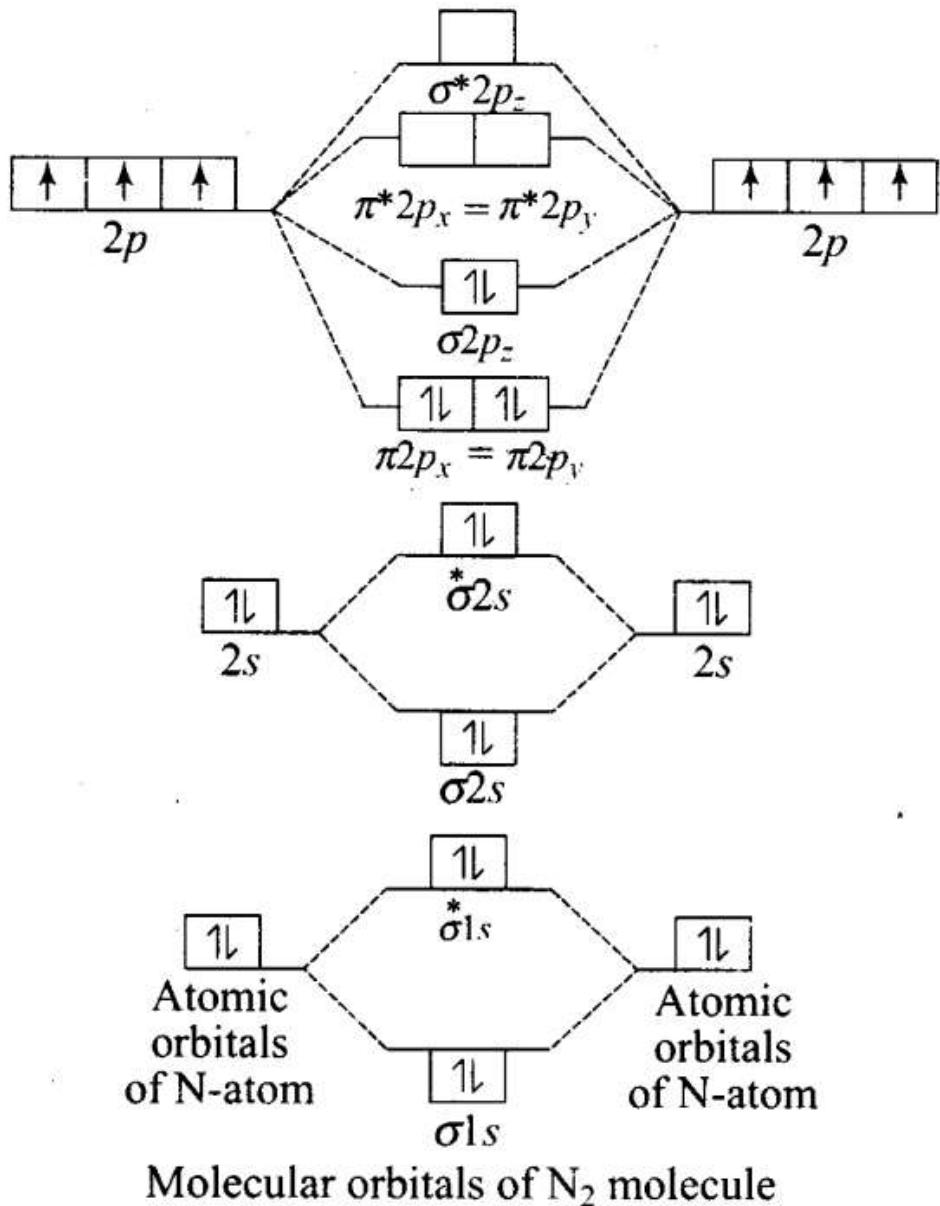
(e) Distinguish between cis- and trans- isomers. Dipole moment measurements help to distinguish between cis- and trans- isomers because ds-isomer has usually higher dipole moment than trans isomer.

(f) Distinguish between ortho, meta and para isomers. Dipole moment measurements help to distinguish between o-, m- and p-isomers because the dipole moment of p-isomer is zero and that of o-isomers is more than that of m-isomer.

(ii) CO_2

$$\overset{\delta-}{O}=\overset{2\delta+}{C}=\overset{\delta-}{O} \qquad \mu = 0$$

NF$_3$

Net dipole moment
$\mu = 0.23$ D

CHCl$_3$

Net dipole moment
$\mu = 1.04$ D
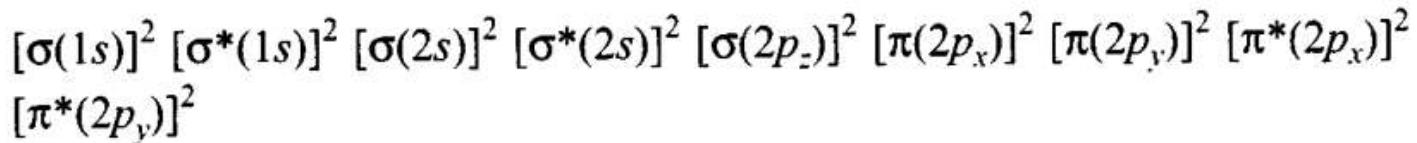
Q61. Use the molecular orbital energy level diagram to show that N_2 would be expected to have a triple bond, F_2 a single bond and Ne_2 no bond.
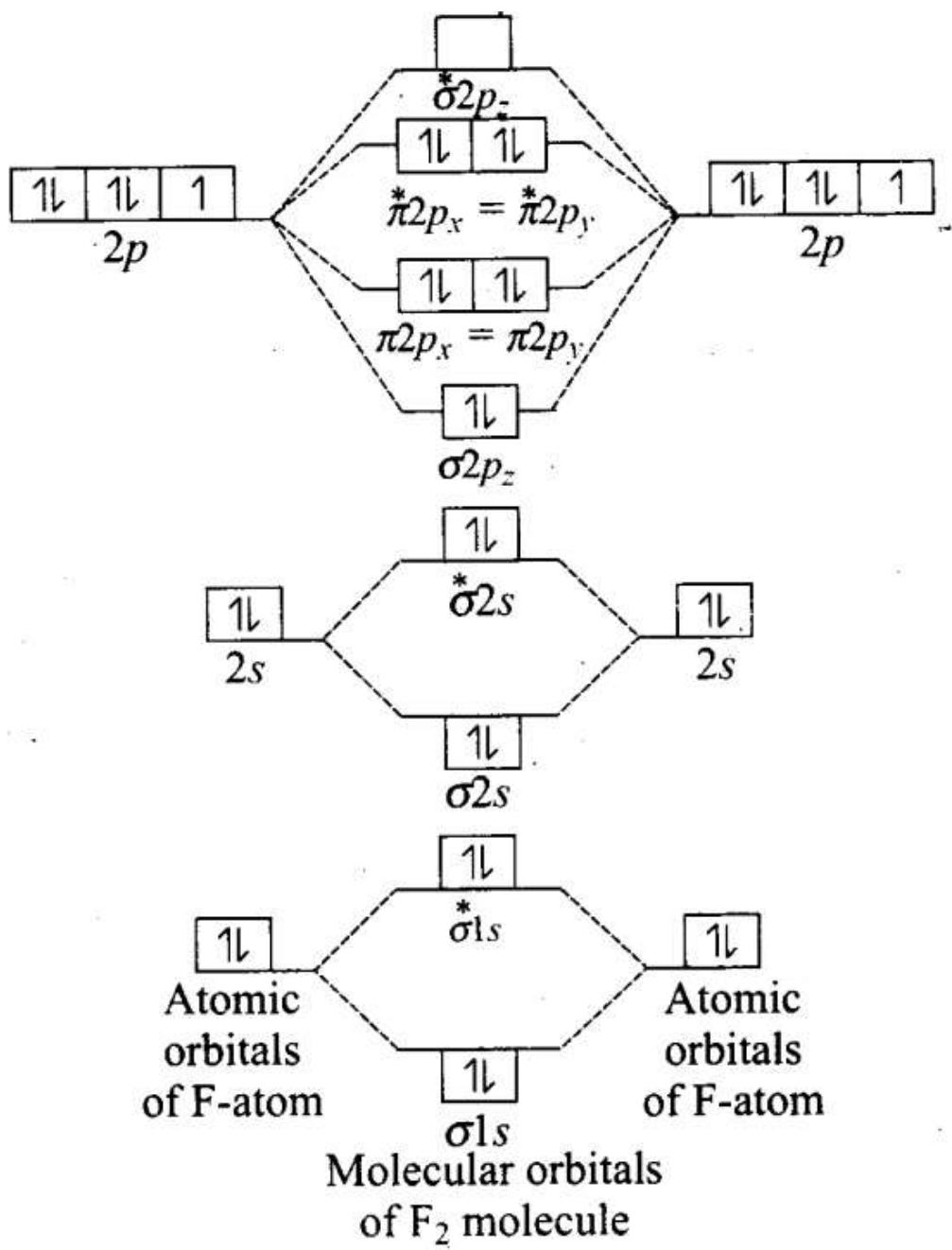
Sol. Nitrogen molecule (N_2): Electron configuration of N-atom ($Z = 7$) is $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$. Total number of electrons present in N_2 molecule is 14, 7 from each N-atom. The electron configuration of N_2 molecule will be

$$\sigma 1s^2 \; \sigma^* 1s^2 \; \sigma 2s^2 \; \sigma^* 2s^2 \; \pi 2p_x^2 \; \pi 2p_y^2 \; \sigma 2p_z^2$$

Molecular orbitals of N_2 molecule

Bond Order $= \dfrac{1}{2}(10 - 4) = 3$ (triple bond)

Fluorine molecule (F_2): Electronic configuration of F-atom ($Z = 9$) is $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^1$. Total number of electron in F_2 molecule is 18, 9 from each F-atom. Thus, it has two electrons more than O_2 molecule. These two additional electrons will enter into $\pi^*(2p_x)$ and $\pi^*(2p_y)$ orbitals, one in each of them so that pairing of electrons takes place in these orbitals. Thus, the electronic configuration of F_2 will be

$$[\sigma(1s)]^2 \, [\sigma^*(1s)]^2 \, [\sigma(2s)]^2 \, [\sigma^*(2s)]^2 \, [\sigma(2p_z)]^2 \, [\pi(2p_x)]^2 \, [\pi(2p_y)]^2 \, [\pi^*(2p_x)]^2$$
$$[\pi^*(2p_y)]^2$$

$\overset{*}{\sigma}2p_z$

$\overset{*}{\pi}2p_x = \overset{*}{\pi}2p_y$

$2p$

$2p$

$\pi2p_x = \pi2p_y$

$\sigma2p_z$

$\overset{*}{\sigma}2s$

$2s$

$2s$

$\sigma2s$

$\overset{*}{\sigma}1s$

Atomic orbitals of F-atom

$\overset{*}{\sigma}1s$

Atomic orbitals of F-atom

$\sigma1s$

Molecular orbitals of F_2 molecule
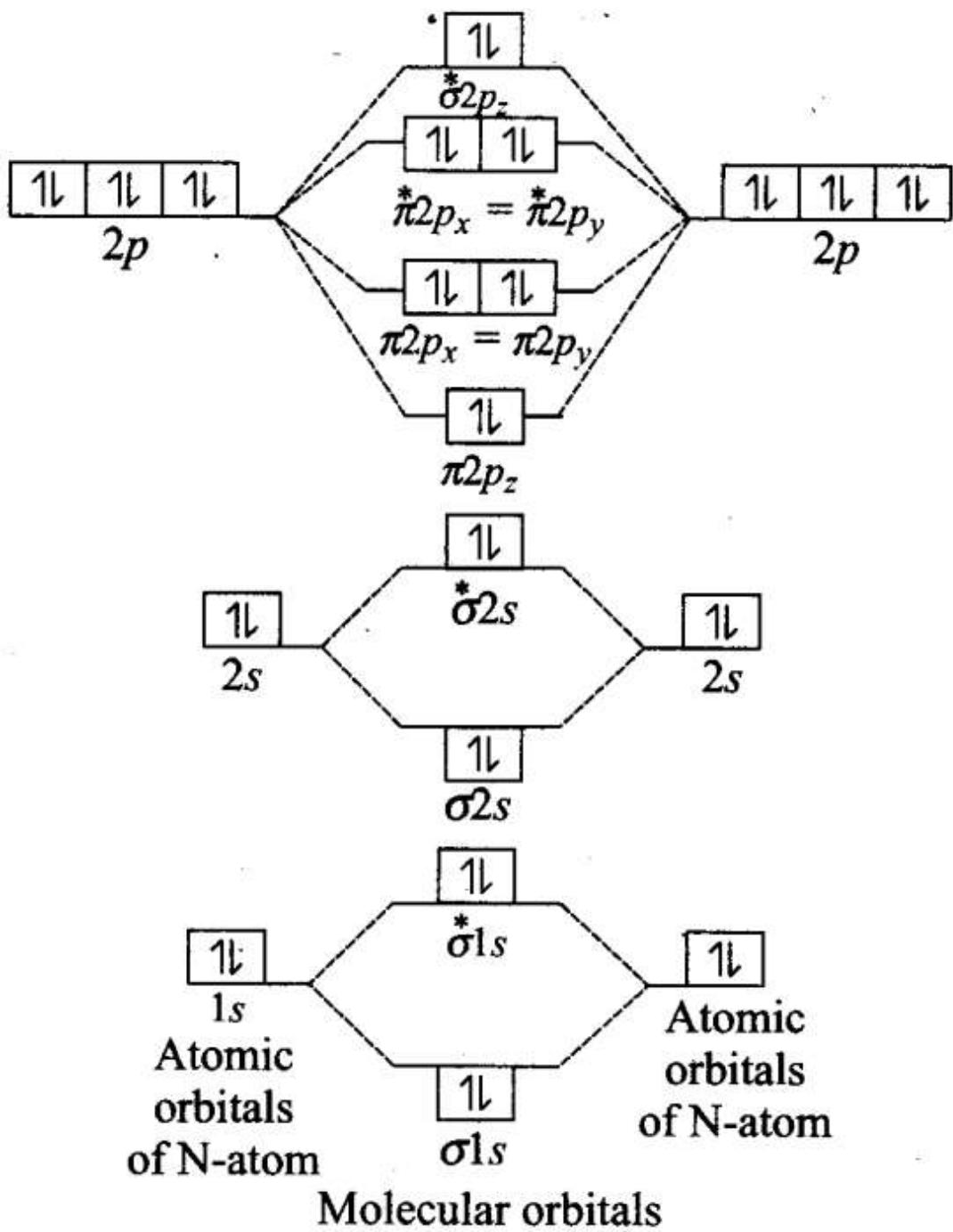
Bond Order = $\dfrac{1}{2}(10-8) = 1$ (single bond)

Ne₂:

Total number of electrons in Ne_2 = 20.

The electron configuration of Ne_2 molecule will be

Ne_2 (20): $\sigma1s^2\ \sigma*1s^2\ \sigma2s^2\ \sigma*2s^2\ \sigma2p_z^2\ \pi2p_x^2 = \pi2p_y^2\ \pi*2p_x^2 = \pi*2p_y^2\ \sigma*2p_z^2$

No bond is formed between two Ne atoms or in other words, Ne_2 does not exist.

Bond Order = ½ (10 − 10) = 0 (No bond)

Q62. Briefly **describe the valence bond theory of covalent bond formation by taking an example of hydrogen. How can you interpret energy changes taking place in the formation of dihydrogen?**

Sol: The valence bond theory was put forward by Heitler and London in 1927. It was later improved and developed by L. Pauling and J.C. Slater in 1931. The valence bond theory is based on the knowledge of atomic orbitals and electronic configurations of elements, overlap criteria of atomic orbitals and stability of molecule.

The main points of valence bond theory are
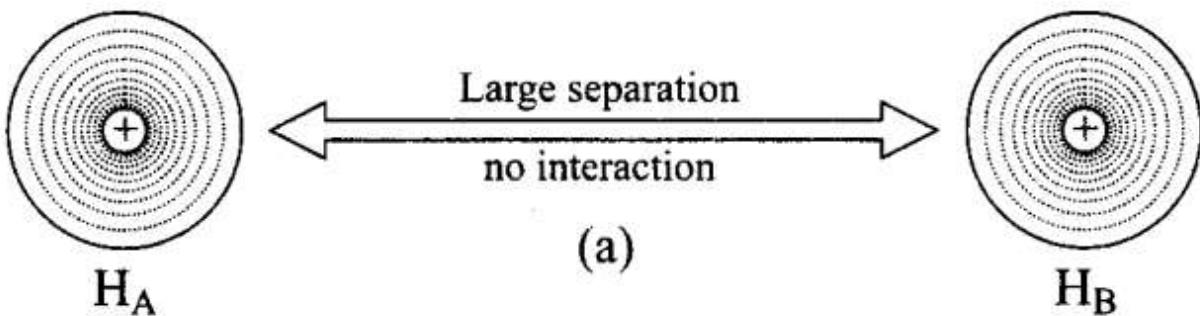
(i) Atoms do not lose their identity even after the formation of the molecule.

(ii) The bond is formed due to the interaction of only the valence electrons as the two atoms come close to each other. The inner electrons do not participate in the bond formation.

(iii) During the formation of bond, only the valence electrons from each bonded atom lose their identity. The other electrons remain unaffected.

(iv) The stability of bond is accounted by the fact that the formation of bond is accompanied by release of energy. The molecule has minimum energy at a certain distance between the atoms known as intemuclear distance. Larger the decrease in energy, stronger will be the bond formed.

Valence bond Treatment of Hydrogen Molecule:
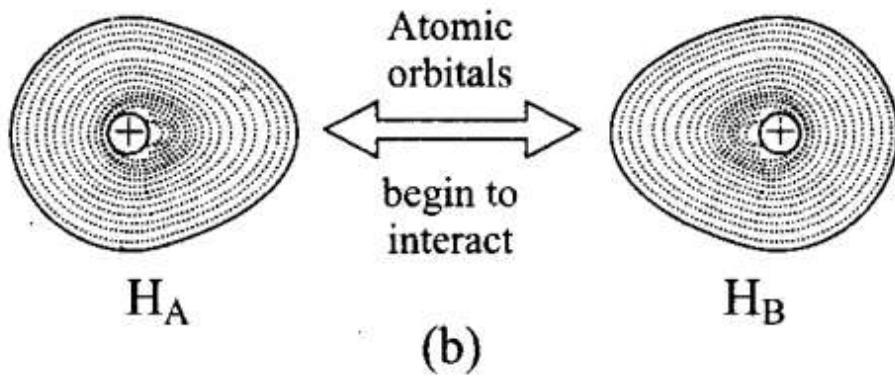
Consider two hydrogen atoms A and B approaching each other havingnuclei H_a and H_B and the corresponding electrons e_A and e_B respectively.

When atoms come closer to form molecules new forces begin to operate.
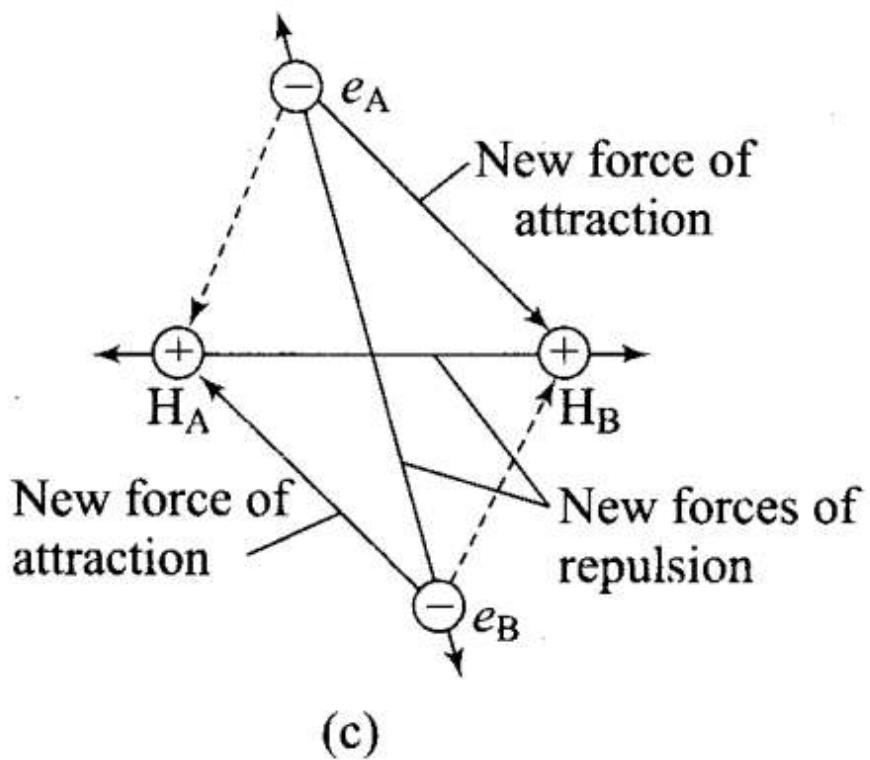
(a) The force of attraction between nucleus of atom and electron of another atom.

(b) The force of repulsion between two nuclei of the atom and electron of two atoms.

Fig. (a) Two hydrogen atoms at a large distance and hence, no interaction, (b) Two hydrogen atom closer to each other atomic orbitals begin to interact, (c) Attractive and repulsive forces in hydrogen atoms when interaction begins. In case of hydrogen: Figure 'a' shows that two hydrogen atoms are at farthest distances and their electron distribution is absolutely symmetrical.

(a) When two hydrogen atom start coming closer to each other, the electron cloud becomes distorted and new attractive and repulsive forces begin to operate as shown in figure 'c'

(b) In figure 'c' dotted lines show attractive forces present in atom already and bold lines show the new

attractive and repulsive forces.

(c) It has been found experimentally that the magnitude of net attractive forces is more than net repulsive forces. Thus stable hydrogen molecule is formed.

Potential energy diagram for formation of hydrogen molecules:

When two hydrogen atoms are at farther distance, there is no force operating between them, when they start coming closer to each other, force of attraction comes into play and their potential energy starts decreasing. As they come closer to each other potential goes on decreasing, but a point is reached, when potential energy acquires minimum value.

Note:

(a) This distance corresponding to this minimum energy value is called the distance of maximum possible approach, i.e. the point which corresponds to minimum energy and maximum stability.

(b) If atoms come further closer than this distance of maximum possible approach, then potential energy starts increasing and force of repulsion comes into play and molecules starts becoming unstable.
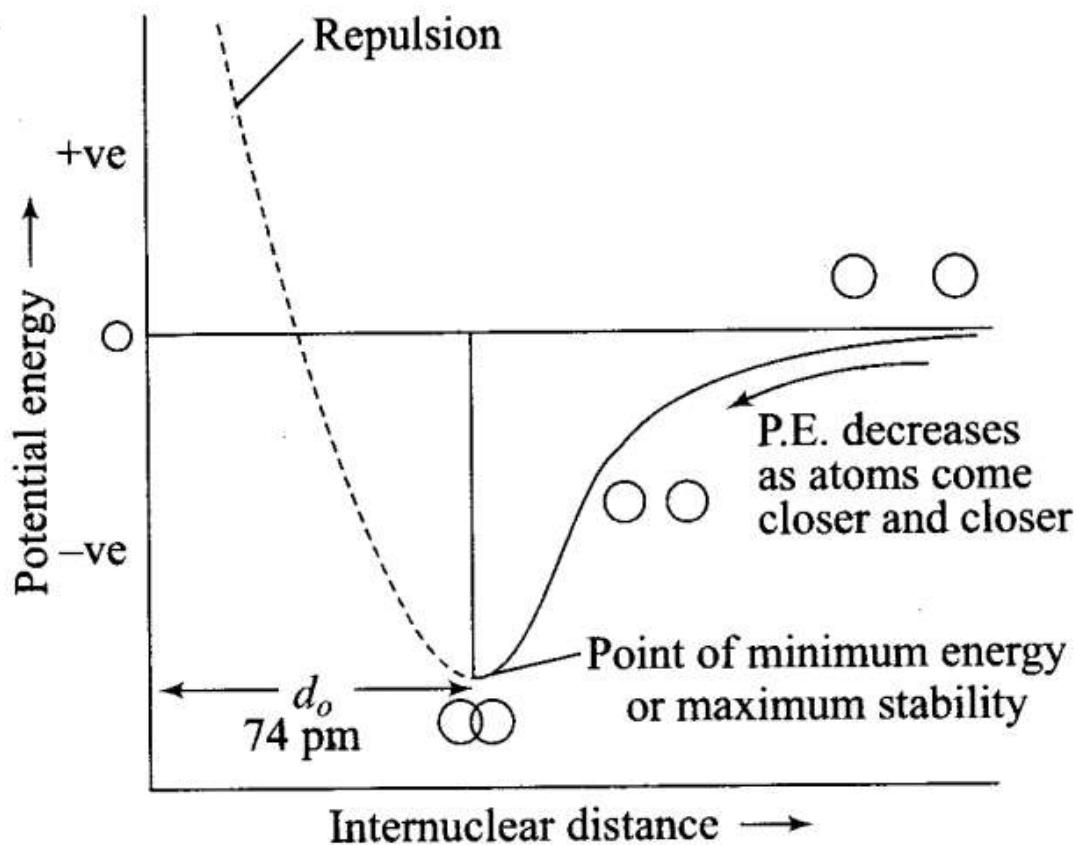

Fig. Variation of potential energy of interaction between two hydrogen atoms