Chemistry Study Materials for Class 11
(NCERT Questions -Answers of Chapter- 04)
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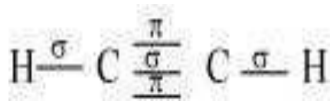
CHEMICAL BONDING AND MOLECULAR STRUCTURE

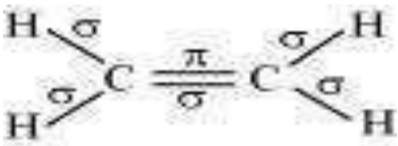
Question 4.28: What is the total number of sigma and pi bonds in the following molecules? (a) C_2H_2 (b) C_2H_4

Answer

A single bond is a result of the axial overlap of bonding orbitals. Hence, it contributes a sigma bond. A multiple bond (double or triple bond) is always formed as a result of the sidewise overlap of orbitals. A pi-bond is always present in it. A triple bond is a combination of two pi-bonds and one sigma bond. Structure of C_2H_2 can be represented as:



Hence, there are three sigma and two pi-bonds in C_2H_2. The structure of C_2H_4 can be represented as:



Hence, there are five sigma bonds and one pi-bond in C_2H_4.

Question 4.29: Considering x-axis as the internuclear axis which out of the following will not form a sigma bond and why?

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

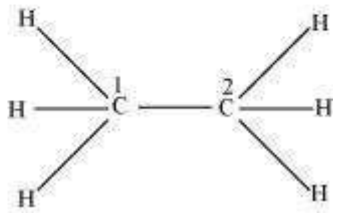
Answer: $2p_y$ and $2p_y$ orbitals will not a form a sigma bond. Taking *x*-axis as the internuclear axis, $2p_y$ and $2p_y$ orbitals will undergo lateral overlapping, thereby forming a pi (π) bond.

Question 4.30: Which hybrid orbitals are used by carbon atoms in the following molecules?

(a)$CH_3–CH_3$; (b) $CH_3–CH=CH_2$; (c) CH_3-CH_2-OH; (d) CH_3-CHO (e) CH_3COOH

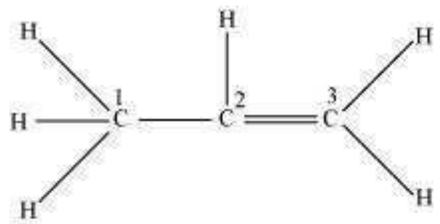
Answer

(a)



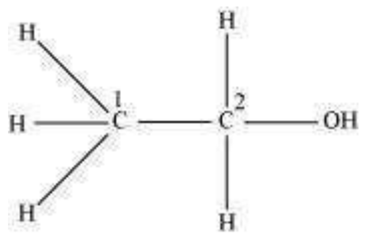
both C_1 and C_2 are sp^3 hybridized.

(b)



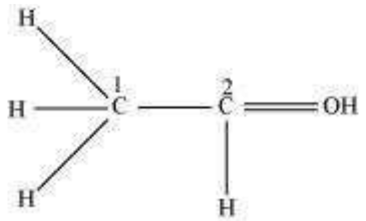
C_1 is sp^3 hybridized, while C_2 and C_3 are sp^2 hybridized.

(c)



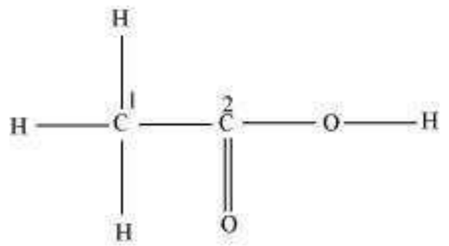
Both C_1 and C_2 are sp^3 hybridized.

(d)



C_1 is sp^3 hybridized and C_2 is sp^2 hybridized.

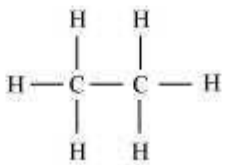
(e) C_1 is sp^3 hybridized and C_2 is sp^2 hybridized.



Question 4.31: What do you understand by bond pairs and lone pairs of electrons? Illustrate by giving one example of each type.

Answer

When two atoms combine by sharing their one or more valence electrons, a covalent bond is formed between them.The shared pairs of electrons present between the bonded atoms are called **bond pairs**. All valence electrons may not participate in bonding. The electron pairs that do not participate in bonding are called **lone pairs** of electrons.



For example, in C_2H_6 (ethane), there are seven bond pairs but no lone pair present.

In H_2O, there are two bond pairs and two lone pairs on the central atom (oxygen).

Question 4.32: Distinguish between a sigma and a pi bond.

Answer: The following are the differences between sigma and pi-bonds:

Sigma (σ) Bond	**Pi (π) Bond**
(a) It is formed by the end to end overlap of orbitals.	It is formed by the lateral overlap of orbitals.
(b) The orbitals involved in the overlapping are *s–s*, *s–p,* or *p–p*.	These bonds are formed by the overlap of *p–p* orbitals only.
(c) It is a strong bond.	It is weak bond.
(d) The electron cloud is symmetrical about the line joining the two nuclei.	The electron cloud is not symmetrical.
(e) It consists of one electron cloud, which is symmetrical about the internuclear axis.	There are two electron clouds lying above and below the plane of the atomic nuclei.
(f) Free rotation about σ bonds is possible.	Rotation is restricted in case of pi-bonds.

Question 4.33: Explain the formation of H_2 molecule on the basis of valence bond theory.

Answer; Let us assume that two hydrogen atoms (A and B) with nuclei (N_A and N_B) and electrons (e_A and e_B) are taken to undergo a reaction to form a hydrogen molecule.

When A and B are at a large distance, there is no interaction between them. As they begin to approach each other, the attractive and repulsive forces start operating.

Attractive force arises between:

(a) Nucleus of one atom and its own electron i.e., $N_A - e_A$ and $N_B - e_B$.

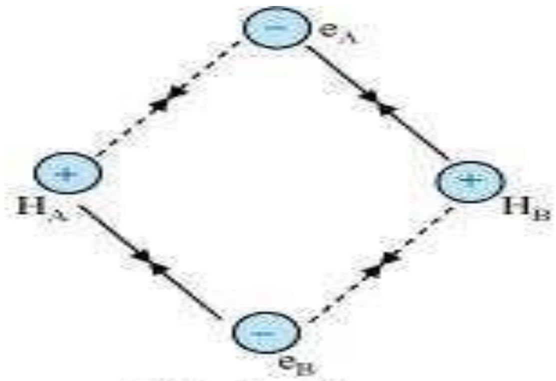
(b) Nucleus of one atom and electron of another atom i.e., $N_A - e_B$ and $N_B - e_A$.

Repulsive force arises between:

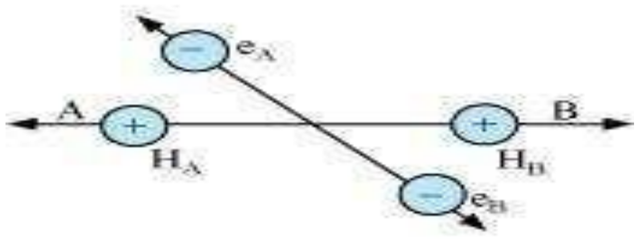
(a) Electrons of two atoms i.e., $e_A - e_B$.

(b) Nuclei of two atoms i.e., $N_A - N_B$.

The force of attraction brings the two atoms together, whereas the force of repulsion tends to push them apart.



Attractive Forces



Repulsive Forces

The magnitude of the attractive forces is more than that of the repulsive forces. Hence, the two atoms approach each other. As a result, the potential energy decreases. Finally, a state is reached when the attractive forces balance the repulsive forces and the system acquires minimum energy. This leads to the formation of a dihydrogen molecule.

Question 4.34: Write the important conditions required for the linear combination of atomic orbitals to form molecular orbitals.

Answer: The given conditions should be satisfied by atomic orbitals to form molecular orbitals:

(a) The combining atomic orbitals must have the same or nearly the same energy. This means that in a homonuclear molecule, the 1*s*-atomic orbital of an atom can combine with the 1*s*-atomic orbital of another atom, and not with the 2*s*-orbital.

(b) The combining atomic orbitals must have proper orientations to ensure that the overlap is maximum

(c) The extent of overlapping should be large.

Question 4.35: Use molecular orbital theory to explain why the Be_2 molecule does not exist.

Answer

The electronic configuration of Beryllium is $1s^2 2s^2$

The molecular orbital electronic configuration for Be_2 molecule can be written as:

$$\sigma_{1s}^2 \; \sigma_{1s}^{*2} \; \sigma_{2s}^2 \; \sigma_{2s}^{*2}$$

Hence, the bond order for Be_2 is $\frac{1}{2}(N_b - N_a)$.

Where,

N_b = Number of electrons in bonding orbitals

N_a = Number of electrons in anti-bonding orbitals

∴ Bond order of Be_2 $= \frac{1}{2}(4-4) = 0$

A negative or zero bond order means that the molecule is unstable.

Hence, Be_2 molecule does not exist.

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