Chapter – 4 (Chemical Bonding and Molecular Structures)

Exercise Questions:

Question :1 Explain the formation of a chemical bond.

Answer:1

A chemical bond is defined as an attractive force that holds the constituent (atoms, ions etc.) together in a chemical species.

Various theories have been suggested for the formation of chemical bonds such as the electronic theory, valence shell electron pair repulsion theory, valence bond theory and molecular orbital theory.

A chemical bond formation is attributed to the tendency of a system to attain stability. It was observed that the inertness of noble gases was because of their full filled outermost orbitals.

Hence, it was postulated that the elements having incomplete outermost she'll are unstable. Atoms therefore combine with each other and complete their respective octet or fillets to attain the stable configuration of the nearest noble gases. This combination can occur either by sharing of electrons or by transferring one or more electrons from one atom to another. The chemical bond formed as a result of sharing of electrons between atoms is called a covalent bond. An ionic bond is formed as a result of transfer of electrons from one atom to another.

Question :2 Write Lewis dot symbol for atoms of the following elements : Mg, Na, B, O, N, Bring.

Answer:

G.N Lewis introduced symbols to represent the valence electrons in the atoms.They are called Lewis symbols, where valence electrons are shown as dots surrounding the symbol of atom or ion.

Mg: Mg belongs to 3rd period & 2nd group of periodic table & has 2 electrons in the valence shell.

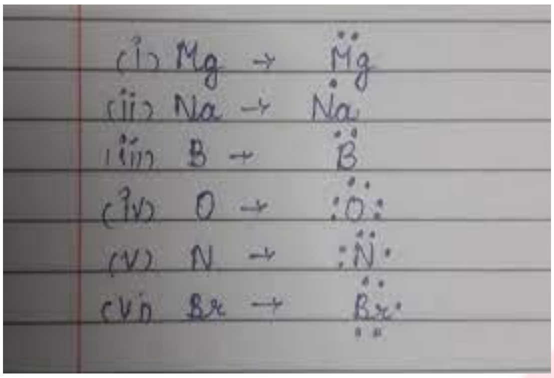
Na: Na has electronic configuration of [Ne]10 3s1, i.e it has 1 electron in outer most shell.

B: Boron belongs to 2nd period & 13th group with 3 electron in valence shell.

O: Oxygen has electronic configuration of [He]2 2s2 2p4,i.e having 6 electrons in valence shell.

N: Nitrogen has electronic configuration of [He]2 2s2 2p3. Hence, the Lewis dot structure is:

Br: There are seven valence electrons in bromine.

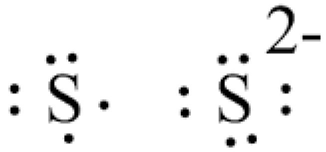

(i) Mg → Mg
(ii) Na → Na
(iii) B → B
(iv) O → O
(v) N → N
(vi) Br → Br

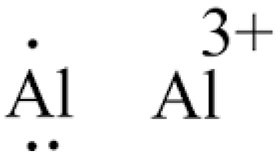
Question :3 Write Lewis symbols for the following atoms and ions:
S and S- ,Al and Al3+, H and H-

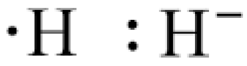
Answer:

(i) S and S2– Electronic configuration of S is [Ne]10 3s23p4, i.e it has 6 electrons in its valence shell,therefore The Lewis dot symbol of sulphur (S) is . The dinegative charge shows that there will be two electrons more in addition to the six valence electrons. The electronic configuration can be [Ne]10 3s^{2}3p^6

(ii) Al and Al3+ The electronic configuration of Aluminium is [Ne]10 3s2 3p1, i.e the valence electron is 3,therefore The Lewis dot symbol of aluminium (Al) is . The tripositive charge on a species infers that it has donated its three electrons to other atom & its possible electronic configuration is [Ne]10 3s1.


(i)

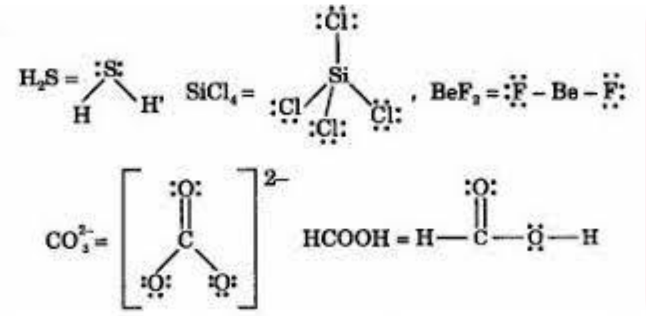

(ii)


(iii)

(iii) H and H– Hydrogen has electronic configuration of 1s1,it has 1 electron in its valence shell. The Lewis dot symbol of hydrogen (H) is . The uninegative charge infers that there will be one electron more in addition to the one valence electron. The possible configuration may be 1s2. Hence, the Lewis dot symbol is

Question :4 Draw the Lewis structures for the following molecules and ions: H2S, SiCl2, BeF2, Co3*2-, HCOOH.

Answer:



Question :5 Define octet rule. Write it's significant and limitations.

Answer:

The octet rule refers to the tendency of atoms to prefer to have eight electrons in the valence shell. When atoms have fewer than eight electrons, they tend to react and form more stable compounds. It was postulated by Kossel & Lewis. The atoms complete their octet by either loosing or gaining of electrons. They do so to attain the nearest noble gas stable electronic configuration .

The octet rule successfully explained the formation of chemical bonds depending upon the nature of the element.

Limitations of the octet theory:

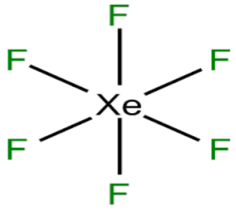
1) The main limitation to the rule is hydrogen, which is at its lowest energy when it has two electrons in its valence shell. Helium (He) is similar in that it, too, only has room for two electrons in its only valence shell.

Hydrogen and helium have only one electron shell. The first shell has only one s orbital and no p orbital, so it holds only two electrons. Therefore, these elements are most stable when they have two electrons.

2) Second limitation are aluminum and boron, which can function well with six valence electrons. Consider BF3. The boron shares its three electrons with three fluorine atoms. The fluorine atoms follow the octet rule, but boron has only six electrons. Most elements to the left of the carbon group have so few valence electrons that they are in the same situation as boron: they are electron deficient.

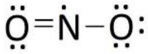
3) In Period 3, the elements on the right side of the periodic table have empty d orbitals. The d orbitals may accept electrons, allowing elements like sulfur and phosphorus to have more than an octet.

Compounds such as PCl_5 and SF_6 can form. These compounds have 10 and 12 electrons around their central atoms, respectively.



Xenon hexafluoride uses d-electrons to form more than an octet. This compound shows another exception: a noble gas compound.

4) The octet rule is not satisfied for all atoms in a molecule having an odd number of electrons. For example, NO_2 do not satisfy the octet rule.



5) The rule failed to predict the shape and relative stability of molecules

Question :6 Write the favourable factors for the formation of ionic bond.

Answer:

An ionic bond is formed by the transfer of one or more electrons from one atom to another. Hence, the formation of ionic bonds depends upon the ease with which neutral atoms can lose or gain electrons. Bond formation also depends upon the lattice energy of the compound formed.

Hence, favourable factors for ionic bond formation are as follows :

I. Low ionization enthalpy of metal atoms.
II. High electron gain enthalpy of a non metal.
III. High lattice energy of the compound formed.

Question :7 Discuss the shape of the following molecules using the VESPR model: $BeCl_2$, BCl_3, $SiCl_2$, AsF_5, H_2S, PH_3.

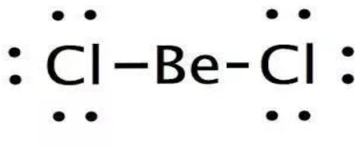
Answer:

Valence shell electron pair repulsion (VSEPR)theory :

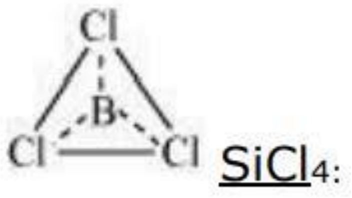
It is a model used to predict the 3D geometry of individual molecules from the number of electron pairs surrounding their central atoms. It was developed by Gillespie & Nyholm.

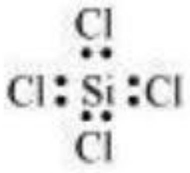
$BeCl_2$:

The central atom has no lone pair and there are two bond pairs. i.e., $BeCl_2$ is of the type EX_2. Hence, it has a linear shape.



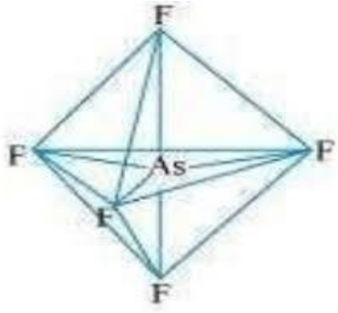
BCl3:
The central atom has no lone pair and there are three bond pairs. Hence, it is of the type EX3. Hence, it is trigonal planar.



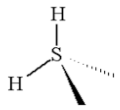


SiCl4:
The central atom has no lone pair and there are four bond pairs. Hence, the shape of SiCl4 is tetrahedral being the EX4 type molecule.

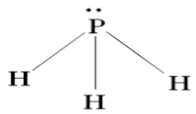
AsF5:
The central atom has no lone pair and there are five bond pairs. Hence, AsF5 is of the type EX5. Therefore, the shape is trigonal bipyramidal.



H2S:
The central atom has one lone pair and there are two bond pairs. Hence, H2S is of the type AB2E. The shape is Bent.

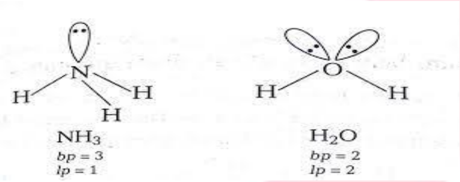


PH3:
The central atom has one lone pair and there are three bond pairs. Hence, PH3 is of the EX5 type. Therefore, the shape is trigonal bipyramidal.



Question :8 Although geometry of NH3 and H2O molecules are distorted tetrahedral, bond angle in water is less than that of ammonia. Discuss.

Answer :
The molecular geometry of NH3 and H2O can be shown as:



The central atom (N) in NH3 has one lone pair and three bond pairs. In H2O there are two lone pairs and two bond pairs.
The two lone pairs present in the oxygen atom of H2O molecule repel the two bond pairs . This repulsion is stronger that the repulsion between the lone pair and the three bond pairs on the nitrogen atom.
Since the repulsion on the bond pairs in H2O molecule are greater than that in NH3, the bond angle in water is less than that of ammonia.

Question :9 How do you express the bond strength in terms of bond order?

Answer:
Bond strength represents the extent of bonding between two atoms forming a molecule. The larger the bond energy, the stronger is the bond and the greater is the bond order.

Question :10 Define the bond length.

Answer :Bond length is defined as the equilibrium distance between the nuclei of two bonded atoms in a molecule.
Bond length are expressed in terms of Angstrom or picometer and are measured by spectroscopic X – ray diffraction and electron-diffraction techniques.

In an ionic compound, the bond length is the sum of their consent radii(d = RA + rB)

Question : 11 Explain the important aspects of resonance with reference to the (CO3)2-ion.

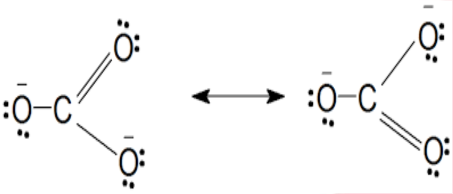
Answer :

According to experimental findings, all carbon to oxygen bonds in CO3)2- are equivalent.

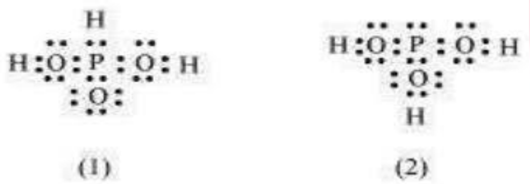
Hence, it is inadequate to represent CO3)2- ion by a single Lewis structure having two single binds and one double bond.

Therefore, carbonate ion is described as a resonance hybrid of the following structures

Image of carbonate ion



Question :12 H3PO3 can be represented by structures 1 and 2 shown below. Can these two structures be taken as the canonical forms of the resonance hybrid representing H3PO3 ? If not, give reasons for the same.



Answer:

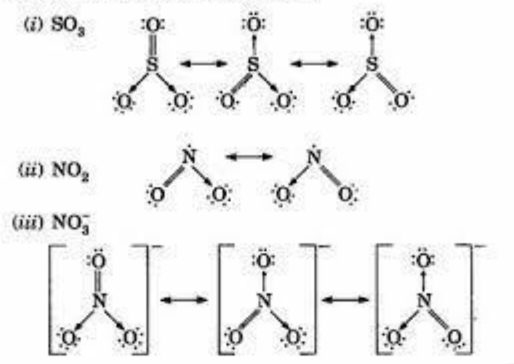
The given structures cannot be taken as the canonical forms of the resonance hybrid of H3PO3 because the positions of the atoms have changed.

Question :13 Write the resonance structures for SO3, NO2 and (NO3)-.

Anawer:

The resonating structures of the following are as:

Resonance structures are as under:



Question :14 Use Lewis symbols to show electrons transfer between the following atoms to form cations and anions :

- **a.) K and S**
- **b. CA and O**
- **c. Al and N**

Answer:

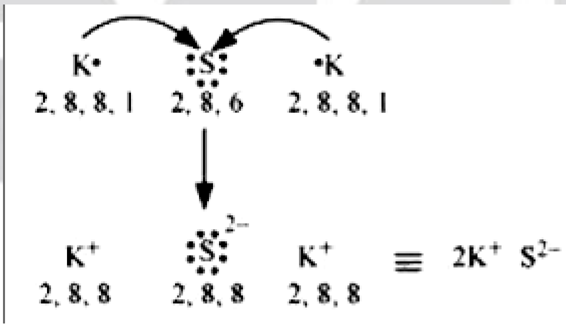
(a) K and S:

The electronic configurations of K and S are as follows:

K: 2, 8, 8, 1

S: 2, 8, 6

Sulphur (S) requires 2 more electrons to complete its octet. Potassium (K) requires one electron more than the nearest noble gas i.e., Argon. Hence, the electron transfer can be shown as:

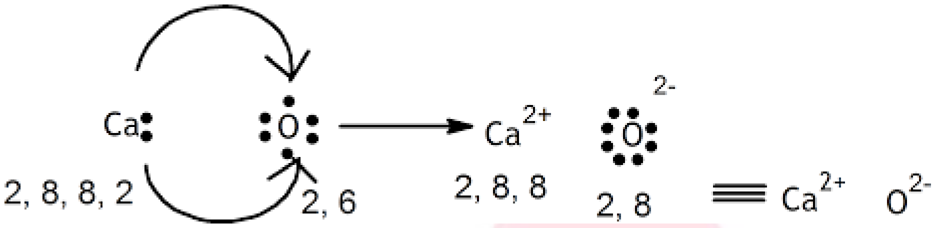


(b) Ca and O: The electronic configurations of Ca and O are as follows:

Ca: 2, 8, 8, 2

O: 2, 6

Oxygen requires two electrons more to complete its octet, whereas calcium has two electrons more than the nearest noble gas i.e., Argon. Hence, the electron transfer takes place as:

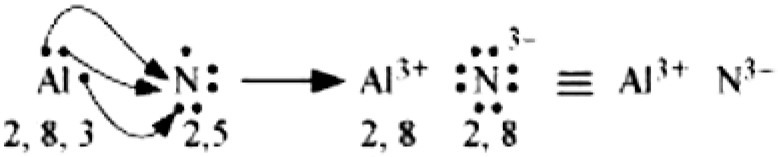


(c) Al and N: The electronic configurations of Al and N are as follows:

Al: 2, 8, 3

N: 2, 5

Nitrogen is three electrons short of the nearest noble gas (Neon), whereas aluminium has three electrons more than Neon. Hence, the electron transference can be shown as



Question : 15 Although both CO2 and H2O are triatomic molecules, the shape of H2O molecules is bent while that of CO2 is linear. Explain this on the basis of dipole moment.

Answer :

According to experimental results, the dipole moment of carbon dioxide is zero. This is possible only if the molecule is linear so that the dipole moment of C-O bonds are equal and opposite to each other.

Resultant = 0D

H2O, on the other hand, has a dipole moment value of 1.84 D. The value of the dipole moment suggests that the structure of H2O molecule is bent where the dipole moment of O-H bonds are unequal.

Question : 16 Write the significance/applications of dipole moment.

Answer:

In heteronuclear molecules, polarization arises due to difference in the electronegativities of the constituents of atoms. As a result, one end of the molecules acquire a positive charge while the other end becomes negative. Hence, a molecule is said to possess a dipole.

The product of the magnitude of the charge and the distance between the centres of positive – negative charges is called the dipole moment of the molecule. It is a vector quantity and is represented by an arrow with its tail at the positive centre and head pointing towards a negative centre.

Dipole moment = change x distance separation.

The SI unit of a dipole moment is esu.

1 esu = 3.335 x 10*-30 cm

Dipole moment is the measure of polarity of a bond. It is used to differentiate between polar and non polar bonds since all non polar molecules have zero dipole moments. It is also helpful in calculating the percentage ionic character of a molecule.

Question : 17 Define electronegativity. How does it differ from electron gain enthalpy?

Answer :

Electronegativity is the ability is an atom in a chemical compound to attract a bond pair of electrons towards itself.

Electronegativity of any given element is not constant. It varies according to the element to which it is bound. It is not a measurable quantity. It is only a relative number.

On the other hand, electron gain enthalpy is the enthalpy change that takes place when an electron is added to a neutral gaseous atom to form an anion. It can be negative or positive depending upon whether the electron is added or removed. An element has a constant value of the electron gain enthalpy that can be measured experimentally.

Question : 18 Explain with the help of suitable example polar covalent bond.

Answer :

When two dissimilar atoms having different electronegativities combine to form a covalent bond, the bond pair of electrons is not shared equally. The bond pair shift towards the nucleus of the atom having greater electronegativities As a result, electron distribution gets distorted and the electron cloud is displaced towards the electronegative atom.

As a result, the electronegative atom becomes slightly negatively charged while the other atom becomes slightly positively charged. Thus, opposite poles are developed in the molecules and thus type of a bond is called a polar covalent bond.

HCl, for example, contains a polar covalent bond. Chlorine atom is more electronegative than hydrogen atom. Hence, the bond pair lies towards chlorine and therefore, it acquires a partial negative charge.

Question : 19 Arrange the bonds in order of increasing ionic character in the molecules :LiF, K2O, N2, SO2 and ClF3.

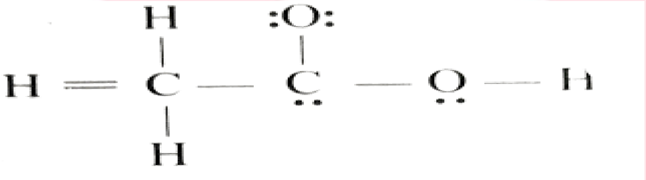
Answer :

The ionic character in a molecule is dependent upon the electronegativity difference between the constituting atoms. The greater the difference, the greater will be the ionic character of the molecule.

On the basis, the order of increasing ionic character in the given molecules is

$N_2 < SO_2 < ClF_3 < K_2O < LiF$

Question :20 The skeletal structure of CH3COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.



Answer:

Concept :- Draw of Lewis structure.

Step 1 :- find the number of valance electrons of all the atoms in the molecule.

Step 2 :- follow the octet rule for all the atoms.

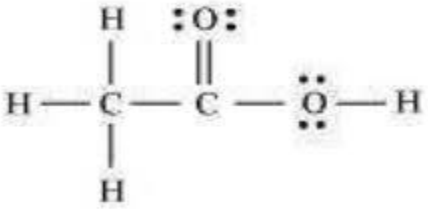
Step 3 :- determine the number of bonds in the molecule, it is find by , number of bonds = (step2 - step1)/2

For example, HF , number of bonds = (10 - 8)/2 = 1

Step 4 :- choose central atom usually, least electronegative atom or , highest valance electrons is central atom.

Step 5 :- now draw by using the number of bonds.

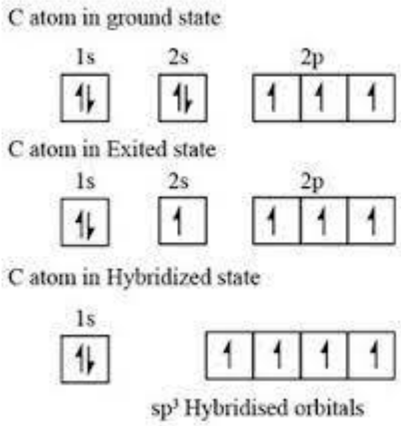
now, you can easily draw the correct Lewis structure.



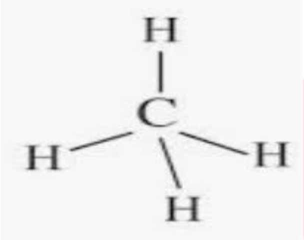
Question : 21 Apartment from tetrahedral geometry, another possible geometry for CH4 is square planar with the four H atoms at the corners of the square and the C atom at its centre. Explain why CH4 is not square planar?

Answer:

Electronic configuration of carbon atom: 6C:1s2 2s2 2p2 In the excited state, the orbital picture of carbon can be represented as:



Hence, carbon atom undergoes sp3 hybridization in CH4 molecule and takes a tetrahedral shape.



For square planar shape, the hybridisation of the central atom has to be dsp*2 .

However, an atom of carbon does not have d – orbitals to undergo dsp*2 hybridisation.

Moreover, with a bond angle of 90 is square planar, the stability of CH4 will be very less because of the repulsion existing between the bond pairs. Hence, VESPER theory also supports a tetrahedral structure for CH4.

Question : 22 Explain why BeH2 molecule has a zero dipole moment although the Be-H bonds are polar.

Answer:

The Lewis structure for BeH2 is as follows :

There is no lone pair at the central atom and there are two bond pairs. Hence, BeH2 is of type AB2 .it has a linear structure.

Dipole moments of each H-Be bond are equal and are in the opposite directions. Therefore, they nullify each other. Hence, BeH2 molecule has zero dipole moment.

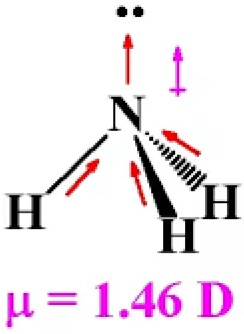
Question : 23 Which out of NH3 and NF3 has higher dipole moment and why?

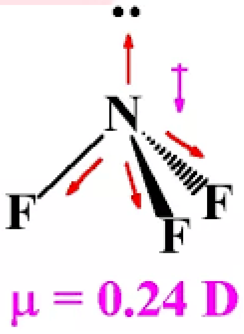
Answer:

In both molecules I. e. NH3 and NF3 ,the central atom has a lone pair electron and there are three bond pairs. Hence, both molecules have a pyramidal shape. Since fluorine is more electronegative than hydrogen, it is expected that the net dipole moment of NF3 is greater than NH3. However, the net dipole moment of NH3 is greater than that of NF3. This can be explained on the basis of directions of the dipole moments of each individual bond in NH3 and NF3. These directions can be shown as:

Thus, the resultant moment of the N-H bond add up to the bond moment of the lone pair, whereas that of the three N-F bonds partly cancels the moment of the lone pair.

Hence, the net dipole moment of NF3 is less than that of NH3.





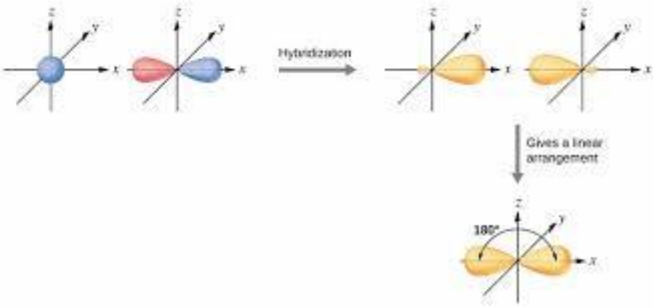
Question :24 What is meant by hybridisation of atomic orbitals? Describe the shapes of so, sp2, sp3 hybrid orbitals.

Answer:

Hybridisation is defined as an intermixing of a set of atomic orbitals of slightly different energies, thereby forming a new set of orbitals having equivalent energies and shapes.

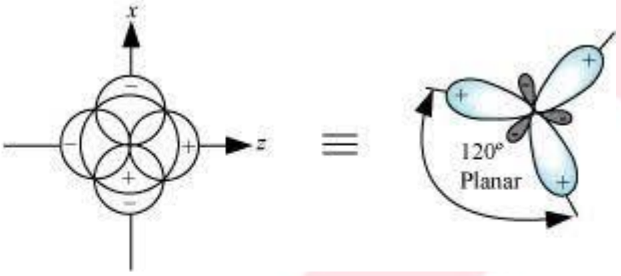
For example, one 2s orbital hybridizes with 2p orbitals of carbon to form three new sp2 hybrid orbitals. These hybrid orbitals have minimum repulsion between their electron pairs and thus, are more stable. Hybridisation helps indicate the geometry of the molecule.

Shape of so hybrid orbitals : so hybrid orbitals have a linear shape. They are formed by the intermixing of s and p orbitals as:



Shape of sp2 hybrid orbitals :

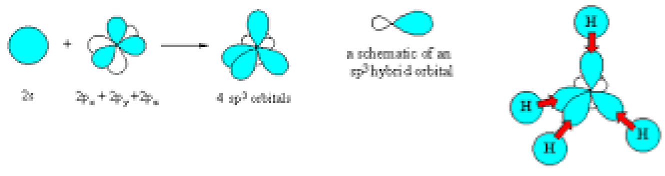
Sp2 hybrid orbitals are formed as a result of the intermixing of one s orbital and two 2p orbitals. The hybrid orbitals are oriented in a trigonal planar arrangement as:



Shape of sp3 hybrid orbitals :

Four sp3 hybrid orbitals are formed by intermixing one s-orbital with three p orbitals

The four sp3 hybrid orbitals are arranged in the form of a tetrahedron as:



sp_3 hybridisation

Question : 25 Describe the change in hybridisation of the AL atom in the following reactions.

$AlCl3 + Cl^- \rightarrow AlCl_4^-$

Answer:

Electronic configuration of Al

In ground state $_{13}Al = 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

In excited state $_{13}Al = 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$

In the formation of AlCl3 ,

Lone pair = (8 × 3 - 3 - 3 × 7)/2 =

Bond pair = 3

Hence, Al undergoes sp^2 hybridisation and it is trigonal planar { because lone pair is zero } in shape .
While in the formation of AlCl4- ,
Lone pair = (4 × 8 - 3 - 4 × 7 -1)/2 = 0
Bond pair = 4
Hence, Al undergoes sp^3 hybridisation and it is in tetrahedral in shape

Question : 26 Is there any change in the hybridisation of B and N atoms as a result of the following reactions?

$BF_3 + NH_3 \rightarrow F_3B.NH_3$

Answer:

The atomic number of boron is 5 & its electronic configuration in ground state is 1s2 2s2 2p1, & in excited state is 1s2 2s1 2p2 , this means its one s & 2 p orbital will take part in hybridization to form sp2 hybrid orbital.
Similarly N atom has atomic number of 7, with electronic configuration of 1s2 2s2 2p3 in its ground state & in it excited state it has sp3 hybridization.
Now according to the question Boron & Nitrogen in reactant stage has sp2 & sp3 hybridization respectively, but on the product side an adduct is formed, wherein Boron has changed it's hybridization to sp3 while Nitrogen remains in same sp3 hybridization

Question : 27 Draw diagrams showing the formation of a double bond and triple bond between carbon atoms in C_2H_4 and C_2H_2 molecules.

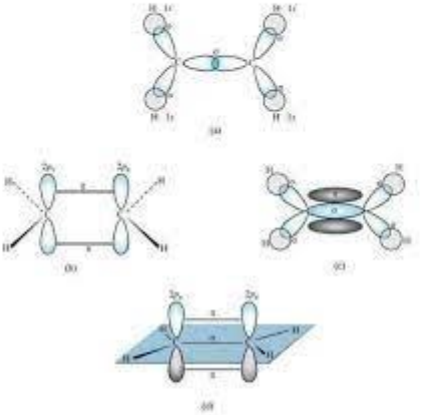
Answer:

C_2H_4:

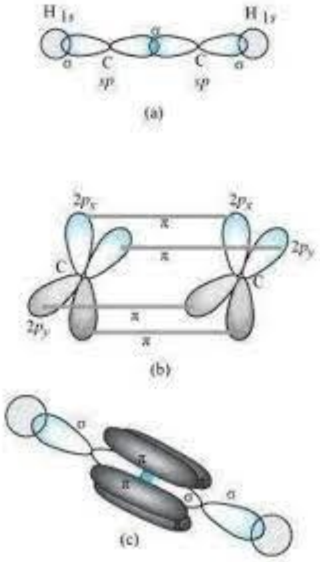
The electronic configuration of C atom in the excited state is:

$C = 1s^2 2s^1 2p_x^1 2p_y^1 2p_z^1$

On the formation of an ethane molecule, one sp2 hybrid orbital of carbon overlaps a sp2 hybridised orbital of another carbon atom, thereby forming a C-C sigma bond. The remaining two sp2 orbitals of each carbon atom from a sp2-s sigma bond with two hydrogen atoms. The unhybridised orbital of one carbon atom undergoes side wise overlap with the orbital of a similar kind present on another carbon atom to form a weak n-bond.



In the formation of C2H2 molecule, each C-atom is so hybridised with two 2p orbitals in an unhybridised state.
One spectrum orbital of each carbon atom overlaps with the other along the interuclear axis forming a C-C sigma bond The second so orbital if each C atom overlaps a half filled 1s orbitals to form sigma bond.



(a)

(b)

(c)

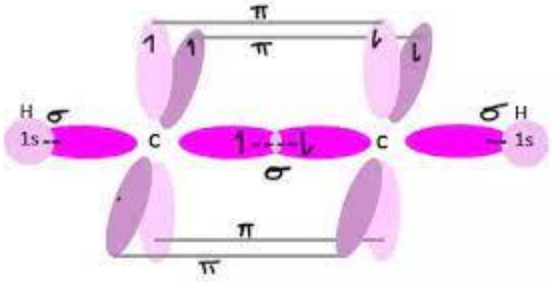
The two unhybridised 2p orbitals of the first carbon undergoes side wise overlap with the 2p orbitals of another carbon atom, thereby forming two pieces bonds between carbon atoms. Hence, the triple bond between two carbon atoms is made up of one sigma and two pieces bonds.

Question : 28 What is the total number of sigma and pick bonds In the following molecules?

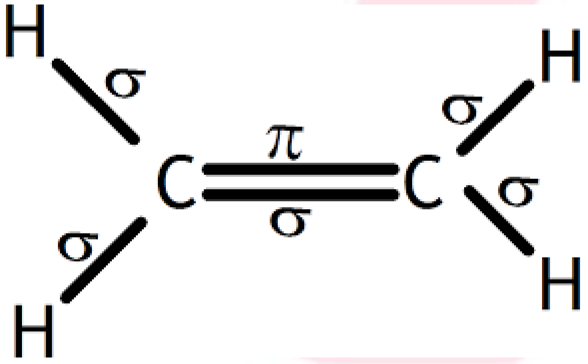
a.) C2H2

b. C2H4

Answer :
A single bond is a result of that axial overlap of bonding orbitals. Hence, it contributes a sigma bond. A multiple born is always formed as a result of the sidewinder overlap of orbitals. A picture bond is always present in it. A triple bond is a contribution of two pieces bonds and one sigma bond.
Structure of C2H2 can be represented as:



Hence, there are three sigma and two pieces bonds in C2H2.
The structure of C2H4 can be represented as:



Hence, there are five sigma bonds and one piece bond in C2H4.

Question : 29 Considering x-axis as the internuclear axis which out of the following will not form a sigma bond and why?

b. 1s and 1s
c. 1s and 2px
d. 2py and 2py
e. 1s and 2s.

Answer:

2py and 2py orbitals will not form sigma bond. Taking x axis as the inter nuclear axis, 2py and 2py orbitals will undergo lateral overlapping, thereby forming a pi bond. Option c.

Question : 30 Which hybrid orbitals are used by carbon atoms in the following molecules?

a. CH3 – CH3
b. CH3 -CH=CH2
c. CH3-CH2-OH
d. CH3-CHO
e. CH3-COOH

Answer:

a.) Both C1 and C2 are sp^3 hybridised.
b.) C1 is sp^3, while C2 and C3 are sp^2 hybridised.
c.) Both C1 and C2 are sp^3 hybridised.
d.) C1 is sp3 and C2 is sp2 hybridised.
e.) C1 is sp3 and C2 is sp2 hybridised.

Question : 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 C2H6 ,there are seven bond pairs but no lone pair present.

In H2O, there are two bond pairs on the central atom.

Question : 32 Explain the formation of H2 molecule on the basis of valence bond theory.

Answer :

Let us assume the two hydrogen atoms with nuclei and electrons are taken to undergo a reaction to form 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 forces arises between :

a. Nucleus of one atom and it's own electron. NA – eA and NB – eB
b. Nucleus of one atom and electron of another atom. Example, NA – eB, NB – eA.

Repulsive forces arises between :

a. Electron of two atoms example., eA – eB
b. Nuclei of two atoms. Example, NA – NB.

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

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. Thus leads to the formation of a dihydrogen molecule.

Question : 33 Write the important conditions required for the linear combination of atomic orbitals to form molecular orbitals.

Answer :

The given condition should be satisfied by atomic orbitals of 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 1s orbital of an atom can combine with the 1s atomic orbital of another atom, not with the 2s 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 : 34 Use molecular orbital theory to explain why the Be2 molecule does not exist.

Answer:

The electronic configuration of Beryllium is 1s2 2s2.

From the electronic configuration it is clear that there is no singly filled atomic orbital present in beryllium.

Without the half filled orbital,the overlapping is not possible ,therefore Be2 molecule does not exist.

Question : 35 Compare the relative stability of the following species and indicate their magnetic properties;

O2, O2)+,O2-(superoxide), O2)2-(peroxide)

Answer:

Concept :- bond order is directly proportional to stability.

Magnetic nature depends paired and unpaired electrons.

If molecule has one or more than unpaired electrons it means molecule is paramagnetic nature.

And if molecule has no unpaired electron{ e.g., all are paired electrons } then, molecule is diamagnetic nature.

Electronic configuration of O2(16 electrons)

$\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, (\pi 2px^2 \approx \pi 2Py^2), (\pi^* 2Px^1 \approx \pi^* 2Py^1)$

Na = 6 , Nb = 10

Now, B.O = 1/2 [10 - 6] = 2

It has two unpaired electrons.so, O2 molecule is paramagnetic.

Electronic configuration of O2 + (15 electrons)

$\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, (\pi 2px^2 \approx \pi 2Py^2), (\pi^* 2Px^1 \approx \pi^* 2Py)$

Na = 5, Nb = 10

Now, B.O = 1/2[10 - 5] = 2.5

It has one unpaired electron so, it is paramagnetic

Electronic configuration of O2-(17 electrons)

$\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, (\pi 2px^2 \approx \pi 2Py^2), (\pi^* 2Px^2 \approx \pi^* 2Py^1)$

Na = 7, Nb = 10

Now, B.O = 1/2[10 - 7] =1.5

It has one unpaired electron so, it is paramagnetic .

Electronic configuration of O2^2-(17 electrons)

$\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, (\pi 2px^2 \approx \pi 2Py^2), (\pi^* 2Px^2 \approx \pi^* 2Py^2)$

Now, B.O = 1/2[10 -8] = 1

It has no unpaired electron.so, it is diamagnetic

Now, bond order is directly proportional to stability so, higher the bond order will be higher stable molecule or ion.

Hence, increasing order of stability is

O2+ > O2 > O2- > O2^2

Question :36 Write the significance of a plus and a minus sign shown in representing the orbitals.

Answer :

Molecular orbitals are represented by wave functions. A plus sign in an orbital indicates a positive wave function while a minus sign in an orbital represents a negative wave function.

Question : 37 Describe the hybridisation in case of PCl_5. Why are the axial bonds longer as compared to equatorial bonds?

Answer:

PCl_5 is sp3d hybridised molecule. It has trigonal bipyramidal geometry. In this case the axial bonds are slightly longer than the equatorial bonds. This is because the axial bonds experience greater repulsion from other bonds than the equatorial bonds.

Question : 38 Define hydrogen bonds. Is it weaker or stronger than the van see Waals forces?

Answer :

A hydrogen bonds is defined as an attractive force acting between the hydrogen attached to an electronegative atom of one molecule and an electronegative atom of different molecule.

Due to difference between electronic negativities, the bond pair between hydrogen and the electronegative atom gets drifted for away from the hydrogen atom. As a result, a hydrogen atom becomes electropsitive with respect to the other atom and acquires a positive charge.

The magnitude of H bonding is maximum in the solid state and minimum in the gaseous state

There are two types of H Bonds;

I. Intermolecular H bond. Example, HF, H_2O etc.

II. Intramolecular H bond. Example, o-nitrophenol.

Hydrogen bonds are stronger than Vander Walls forces since hydrogen bonds are regarded as an extreme form of dipole – dipole interaction.

Question : 39 What is meant by the term bond order? Calculate the bond order of N_2, O_2, O_2^+, O_2^-.

Answer :

Bond order is defined as one half of the difference between the number of electrons present in the bonding and anti bonding orbitals of a molecule.

If N_a is equal to the number of electrons in an anti bonding orbital, then N_b is equal to the number of electrons in a bonding orbital.