Q-1) Explain the formation of a chemical bond

Ans.)

“Chemical bond is an attractive force that bounds the constituents of a chemical species together.”

So many theories are suggested for chemical bond formation such as valence shell electron pair repulsion theory, electronic theory, molecular orbital theory and valence bond theory.

Formation of a chemical bond is credited to the tendency of a system to achieve stability. It was noticed that inertness of noble gasses is the direct result of their completely filled outermost orbitals. Consequently, it was proposed that the elements having a deficiency of electrons in outermost shells are unstable. Thus, atoms combine with one another and finish their separate octets or duplets to achieve the stable configuration of the closest inert gasses. So, this combination may occur either by sharing of electrons. The formed chemical bond is a result of sharing of electrons among atoms is known as a covalent bond. Also, a formed ionic bond is a result of sharing of electrons among atoms.

Q-2) Write Lewis dot symbols for atoms of the following elements:

a) Mg

b) Na

c) B

d) O

e) N

f) Br

Ans.)

a) Mg

Magnesium atom contains only 2 valence electrons. Thus, the lewis dot symbols for Mg is

b) Na

Sodium atom contains only 1 valence electron. Thus, the lewis dot symbols for Na is $Na\cdot$

c) B

Boron atom contains only 3 valence electrons. Thus, the lewis dot symbols for B is

d) O

Oxygen atom contains only 6 valence electrons. Thus, the lewis dot symbols for O is

e) N

Nitrogen atom contains only 5 valence electrons. Thus, the lewis dot symbols for N is

f) Br

Bromine atom contains only 7 valence electrons. Thus, the lewis dot symbols for Br is

Q-3) Write Lewis symbols for the following atoms and ions:

S and S^{2-}; Al and Al^{3+}; H and H^{-}

Ans.)

For S and S^{2-}

A sulphur atom contains only 6 valence electrons. Thus, the lewis dot symbols for S is

The bi-negative charge on sulphur indicates that it has gained 2 electrons. So, six valance electron plus two gained an electron.

Thus lewis dot symbol is

For Al and Al^{3+}

An aluminium atom contains only 3 valence electrons. Thus, the lewis dot symbols for Al is

The tri-positive charge on aluminium indicates that it has donated 3 electrons.

Thus lewis dot symbol is

$[Al]^{3+}$

For H and H^{-}

A hydrogen atom contains only 1 valence electrons. Thus, the lewis dot symbols for H is

$H\cdot$

The single negative charge on hydrogen indicates that it has gained 1 electron. So, one valance electron plus one gained an electron.

Thus lewis dot symbol is

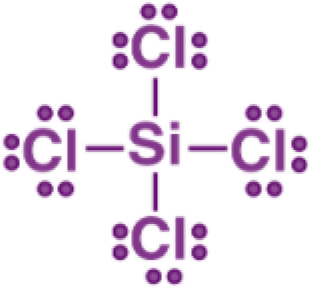
Q-4) Draw the Lewis structures for the following molecules and ions:

H_2S, $SiCl_4$, BeF_2, CO_3^{2-} , HCOOH

Ans.)

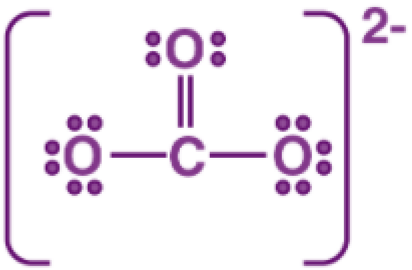
H_2S

$SiCl_4$

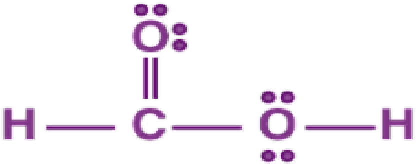


BeF_2

CO_3^{2-}



HCOOH



Q-5) Define the octet rule. Write its significance and limitations
Ans.)

Octet rule says, “atoms can combine either by transfer of valence electrons from one atom to another or by sharing their valence electrons in order to achieve the nearest inert gas configuration by having an octet in their valence shell.”

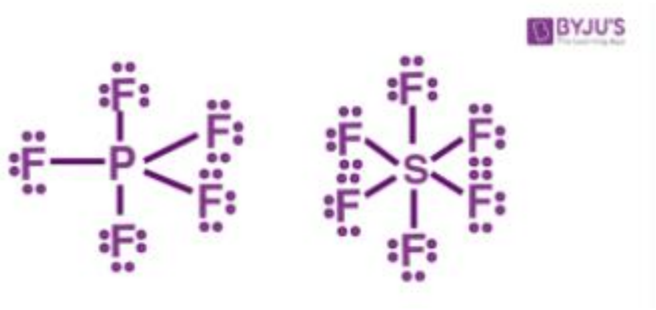
Octet rule explains chemical bond formation depending upon nature of the element.

Limitations:

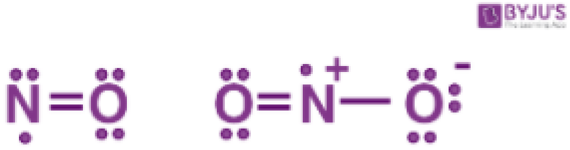
(a) Octet rule fails to predict the relative stability and shape of the molecules.

(b) It is based on the inert nature of noble gases. But, some inert gases say, krypton (Kr) and xenon (Xe) form compounds like KrF_2, XeF_2 etc.

(c) For elements beyond the 3rd period, the octet rule cannot be applied. Elements present beyond 3rd period have more than 8 valence electrons surrounding the central atom. E.g. SF_6, PF_6 etc.



(d) For atoms in a molecule having an odd number of electrons, octet rule is not applied. E.g. For No2 and NO octet rule is not applicable.



(e) If a compound is having less than 8 electrons surrounding the central atom than octet rule cannot be applied to that compound. E.g. BeH_2, $AlCl_3$, LiCl etc. is not obeying octet rule.

Q-6) Write the favourable factors for the formation of an ionic bond.

Ans.)

Formation of ionic bond takes place by transfer of 1 or more electrons from one atom to another. Thus, ionic bond formation depends on the flexibility of neutral atoms to lose or gain electrons. Formation of ionic bond also depends on the lattice energy of the compound which is formed.

The factors that are favourable for ionic bond formation:

(a) High electron affinity of atoms of non- metal.

(b) The high lattice energy of compound which is formed.

(c) Low ionization enthalpy of an atom of metal.

Q-7) Discuss the shape of the following molecules using the VSEPR model:

$BeCl_2$, BCl_3, $SiCl_4$, AsF_5, H_2S, PH_3

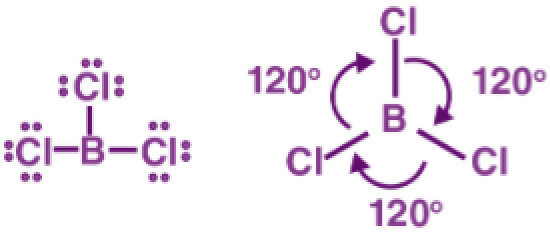
Ans.)

$BeCl_2$

Central atom is not having any lone pair but have 2 bond pairs. Thus, its shape is AB_2. i.e. Linear shape.

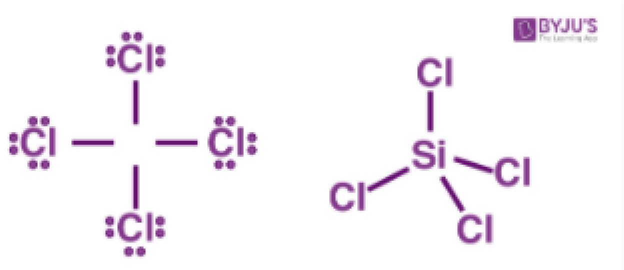
BCl_3

Central atom is not having any lone pair but have 3 bond pairs. Thus, its shape is AB_3. i.e. Trigonal planar.



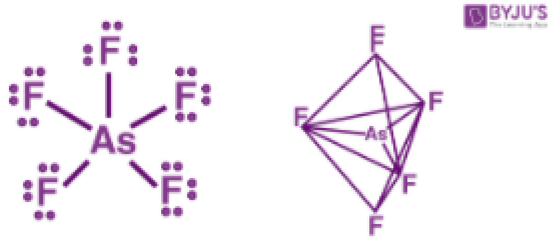
$SiCl_4$

Central atom is not having any lone pair but have 4 bond pairs. Thus, its shape is AB_4. i.e. Tetrahedral.



AsF_5

Central atom is not having any lone pair but have 5 bond pairs. Thus, its shape is AB_5. i.e. Trigonal bipyramidal.

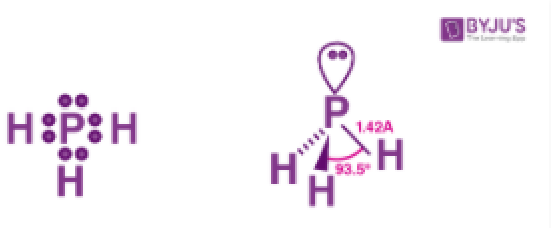


H_2S

Central atom is having 1 lone pair and is having 2 bond pairs. Thus, its shape is AB_2E. i.e. Bent shape.

PH_3

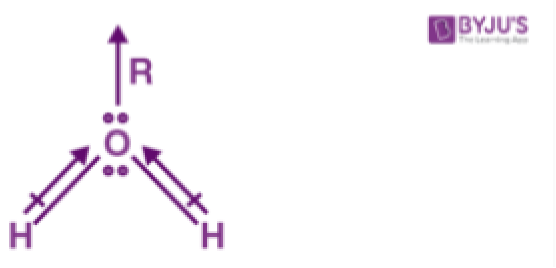
Central atom is having 1 lone pair and is having 3 bond pairs. Thus, its shape is AB_3E. i.e. Trigonal bipyramidal.

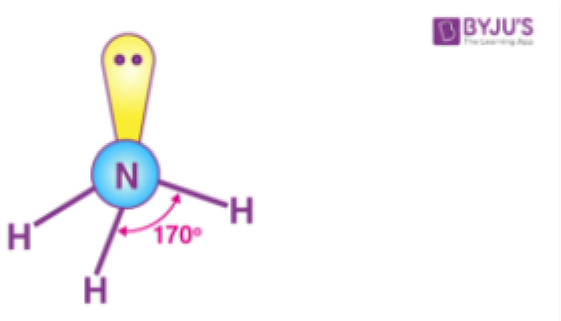


Q-8) Although geometries of NH_3 and H_2O molecules are distorted tetrahedral, bond angle in water is less than that of Ammonia. Discuss.

Ans.)

The geometry of H_2O and NH_3:





Central atom(N) in ammonia is having 1 lone pair and is having 3 bond pairs.

Central atom(O) in water is having 2 lone pair and is having 2 bond pairs.

Thus, these 2 lone pairs on O- atom in water molecule repels the 2 bond pairs. And this repulsion is between lone pair and bond pair on O- atom of H_2O is stronger than the repulsion is between lone pair and bond pair on N-atom of NH_3.

Thus, the bond angle in H_2O is less than NH_3, even though they are having distorted tetrahedral structure.

Q-9) How do you express the bond strength in terms of bond order?

Ans.)

The extent of bonding which occurs between two atoms while forming a molecule is represented by bond strength. As the bond strength increases the bond becomes stronger and the bond order increases.

Q-10) Define Bond length.

Ans.)

“Bond length is defined as the equilibrium distance between the nuclei of 2 bonded atoms in a molecule.”

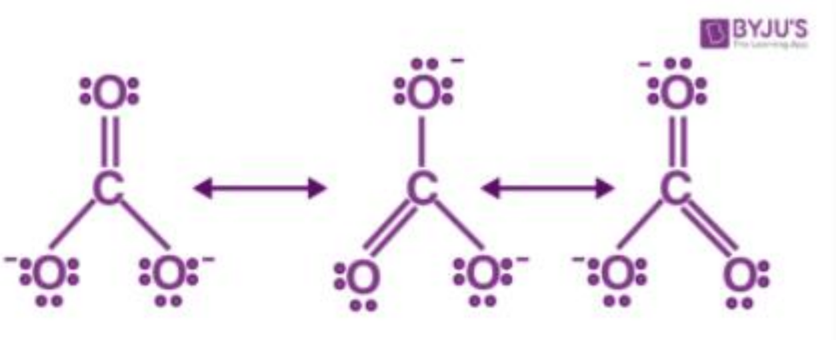
Q-11) Explain the important aspects of resonance with reference to the CO_3^{2-} ion.

Ans.)

Experimental results shows that, all the C-O bond in CO_3^{2-} are equivalent.

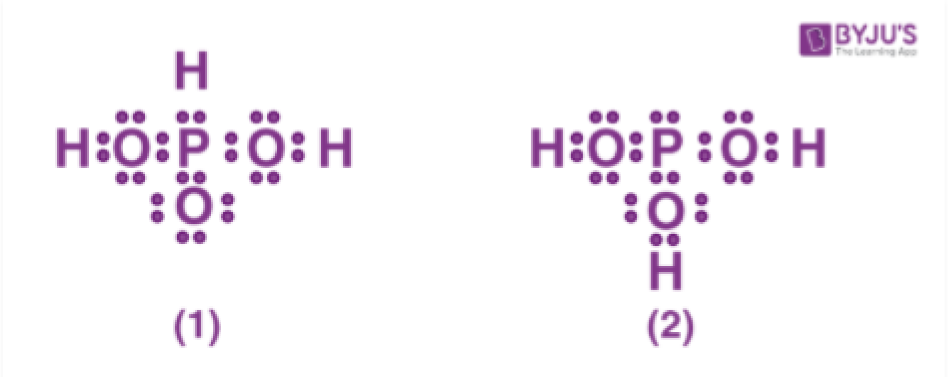
Thus, it is inefficient to represent CO_3^{2-} ion by single lewis structure which is having 1 double bond and 2 single bonds.

Thus, the resonance structures of CO_3^{2-} is :



Q-12) H_3PO_3 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 H_3PO_3? If not, give reasons for the same.

Ans.)

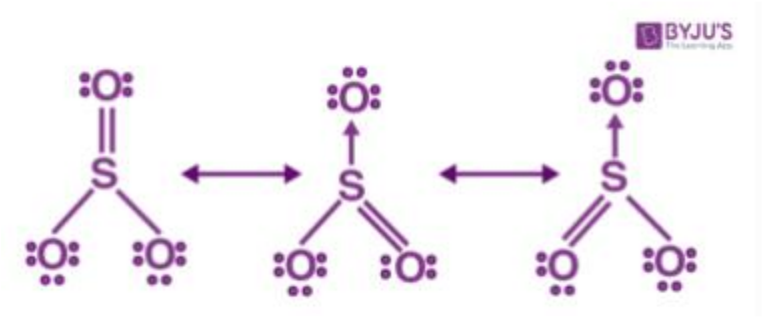


In the given structures the position of atoms are changed, so we cannot take the 2 given structures as a canonical form of resonance hybrid which is representing H_3PO_3.

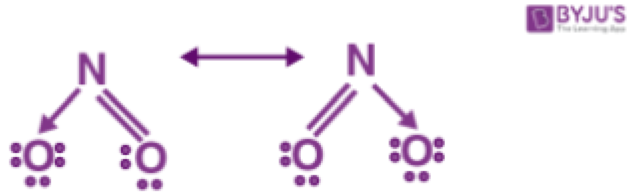
Q-13) Write the resonance structures for SO_3, NO_2, and NO_3^-.

Ans.)

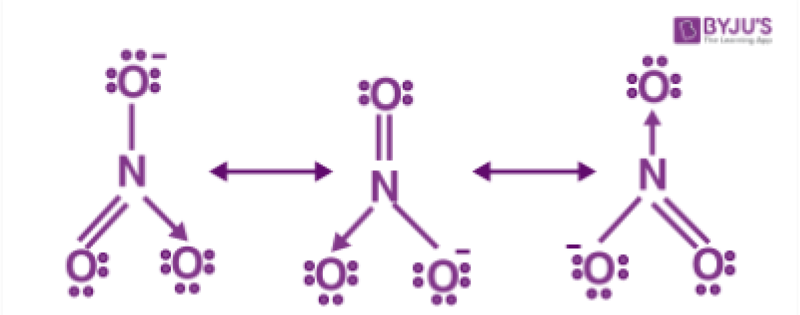
SO_3



NO_2



NO_3^-



Q-14) Use Lewis symbols to show electron transfer between the following atoms to form cations and anions:

(i) K and S

(ii) Ca and O

(iii) Al and N.

Ans.)

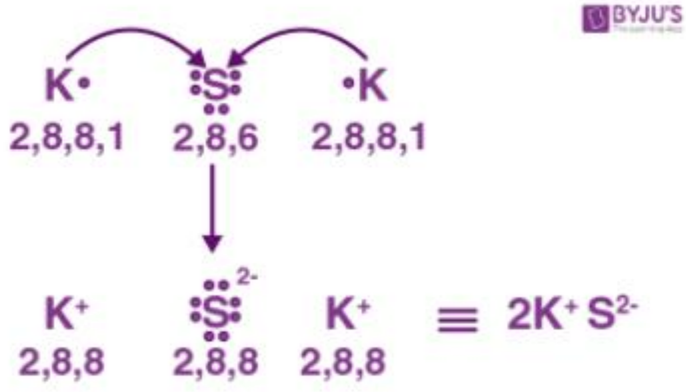
(i) K and S

Electronic configurations of S and K are:

S: 2, 8, 6

K: 2, 8, 8, 1

Here, it clear that K has 1 more electron then nearest inert gas. i.e. Ne, whereas S needs 2 electrons to complete its octet. Thus, the transfer of electrons takes place in the following way,



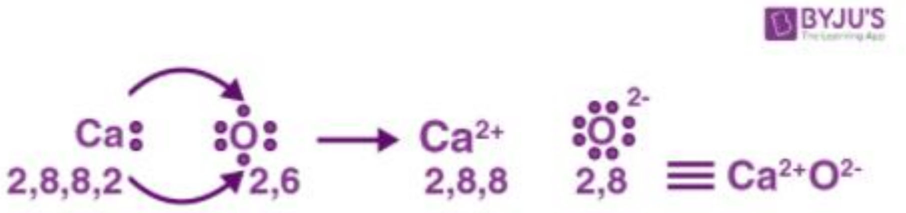
(ii) Ca and O

Electronic configurations of O and Ca are:

O: 2, 6

Ca: 2, 8, 8, 2

Here, it clear that Ca has 2 more electrons then nearest inert gas. i.e. Ar, whereas O needs 2 electrons to complete its octet. Thus, the transfer of electrons takes place in the following way,



(iii) Al and N

Electronic configurations of N and Al are:

N: 2, 5

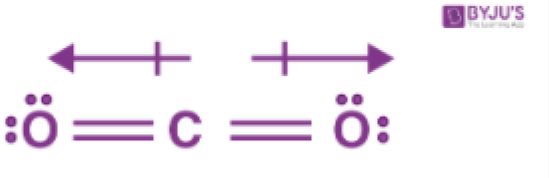
Al: 2, 8, 3

Here, it clear that Al has 3 more electrons then nearest inert gas. i.e. Ne, whereas N needs 3 electrons to complete its octet. Thus, the transfer of electrons takes place in the following way,

Q-15) Although both CO_2 and H_2O are triatomic molecules, the shape of the H_2O molecule is bent while that of CO_2 is linear. Explain this on the basis of dipole moment.

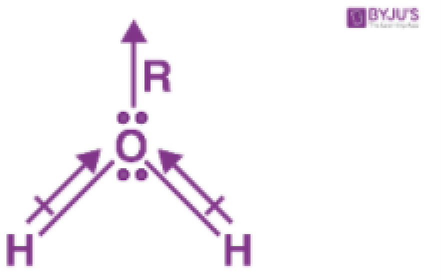
Ans.)

Experimental results show that the dipole moment of CO_2 is 0. And it is possible only if the shape of the molecule is linear as dipole moments of a bond between C-O is equal and opposite so, it nullifies each other.



$$\therefore\ Resultant,\ \mu = 0$$

H_2O has 1.84D dipole moment. The value of dipole moments indicates that the structure of water molecule is bent as dipole moments of the bond between O-H is unequal.



Q-16) Write the significance/applications of dipole moment.

Ans.)

In heteronuclear molecule, there is a difference in electro-negativities of constituents of the atom, due to which polarisation arises it. Thus, one end of acquires positive charge and other end acquires a negative charge. So, molecules are said to have a dipole moment.

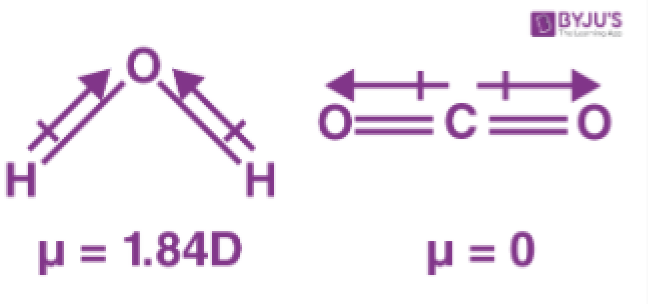
The product of distance between centres of negative and positive charges and magnitude of the charge is known as dipole moment (μ). The dipole moment is a vector quantity and it is represented by an arrow in such a way that the head of a tail is pointing towards negative centre and tail of the arrow is pointing towards the positive centre.

μ = distance between the centres (r) * charge (Q)

SI unit of dipole moment is 'esu'.

1 esu = $3.335 \times 10^{-30} cm$

It is the measure of the polarity of a bond. It differentiates the non-polar and polar bonds as all the non-polar molecules have dipole moment equal to 0. It also calculates the ionic character of a molecule.



Q-17) Define electronegativity. How does it differ from electron gain enthalpy?

Ans.)

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

Sr. No	Electronegativity	Electron affinity
1	A tendency to attract the shared pairs of electrons for an atom which is in chemical compound is its electronegativity.	A tendency to gain electrons for an isolated gaseous atom is its electron gain enthalpy.
2	It varies according to element to which it is bounded with.	It does not vary according to element to which it is bounded with.
3	It is not constant for any element.	It is constant for an element.
4	It is not a measurable quantity.	It is a measurable quantity.

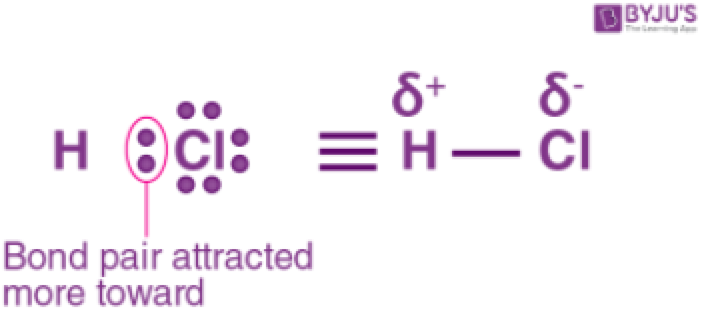
Q-18) Explain with the help of suitable example polar covalent bond.

Ans.)

When two unique atoms having distinct electronegativities join to form a covalent bond, the bond pair of electrons are not shared equally. The nucleus of an atom having greater electro-negativity attracts the bond pair. So, the electron distribution gets distorted and an electronegativity atom attracts the electron cloud.

Thus, the electronegative element gets slightly negatively charged and on the other hand, the other atom gets slightly positively charged. As a result of this, two opposite poles are developed in a molecule and this type of bond formed is termed as 'polar covalent bond'.

E.g. HCl is having a polar covalent bond. In HCl, Cl- atom is having more electronegativity than H- atom. Thus, bond pair shifts towards Cl- atom and because of that, it acquires a positive charge.



Q-19) Arrange the bonds in order of increasing ionic character in the molecules: LiF, K_2O, N_2, SO_2, and ClF_3.

Ans.)

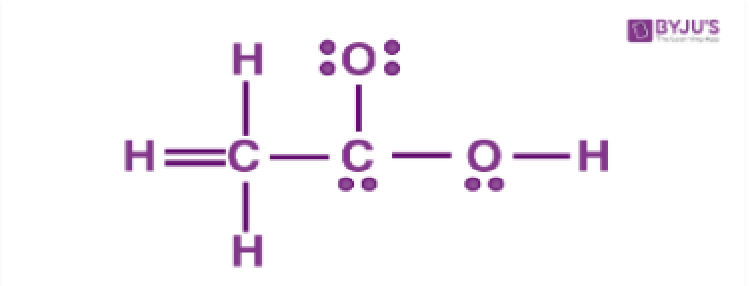
Ionic characteristic of a molecule depends on the difference in electronegativity between constituents atoms. So, higher the difference, the ionic character of a molecule will be higher.

So, the required order of ionic character of the given molecules is

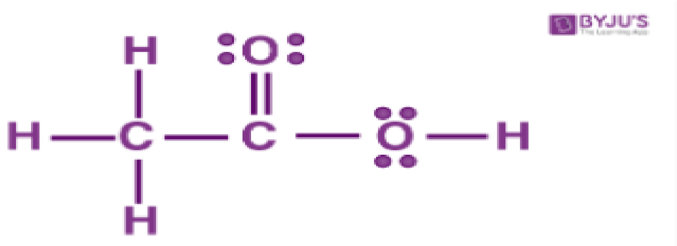
$N_2 < SO_2 < ClF_3 < K_2O < LiF$.

Q-20) The skeletal structure of CH_3COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.

Ans.)



Correct Lewis structure of CH_3COOH is given below:



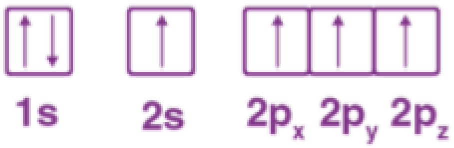
Q-21) Apart from tetrahedral geometry, another possible geometry for CH_4 is square planar with the four H atoms at the corners of the square and the C atom at its centre. Explain why CH_4 is not square planar.

Ans.)

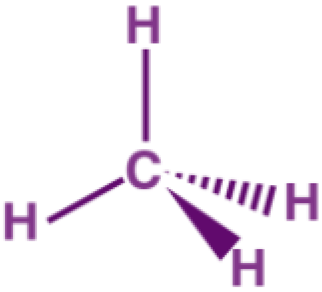
Electronic configuration of C- atom:

$6C : 1s^2\ 2s^2\ 2p^2$

Orbital picture of C- atom in excited state is:



Thus, C- atom undergoes sp^3 hybridization in methane molecule and forms tetrahedral structure.



For square planer geometry, C-atom should have dsp^2 hybridization. But as C- atom is not having d- orbital so it cannot undergo dsp^2 hybridization. Thus, Methane's geometry cannot be square planer.

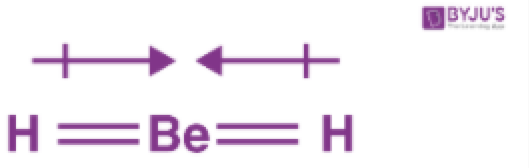
Also in square planar geometry the bond angle is 90° so the stability is not there because of repulsion between bond pairs. So, as per VSEPR theory methane's tetrahedral structure is perfect.

Q-22) Explain why BeH_2 molecule has a zero dipole moment although the Be–H bonds are polar.

Ans.)

Lewis structure of BeH_2 is:

Central atom is not having any lone pair but has 2 bond pairs. Thus, its shape is AB_2. i.e. Linear shape.



Thus, the dipole moment of Be- H bond is equal and opposite in direction so it nullifies one another. Thus, the dipole moment of BeH_2 is 0.

Q-23) Which out of NH_3 and NF_3 has higher dipole moment and why?

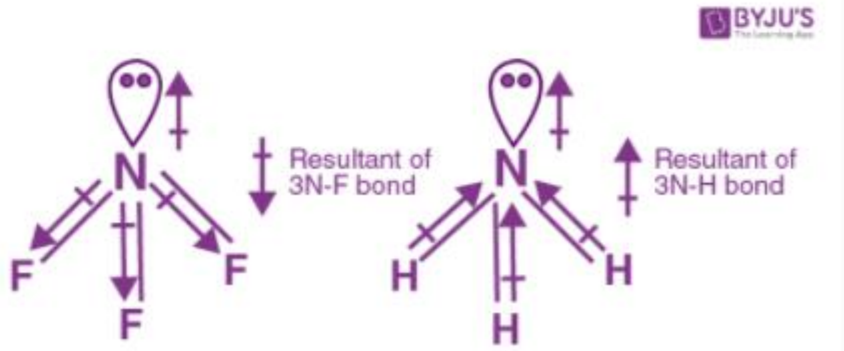
Ans.)

N- atom is the central atom of NF_3 and NH_3.

Central atom is having 1 lone pair and is having 3 bond pairs. Thus, for both the shape is AB_3E. i.e. Pyramidal.

As, F-atom is more electronegativity than H- atom, NF_3 should have higher dipole moment than NH_3. But the dipole moment of NH_3 is 1.46D which is higher than dipole moment of NF_3 which is 0.24D.

It gets clear from the directions of dipole moments of individual bond in NF_3 and NH_3.



As, both the N-H bond are in same direction it adds to the bond moment of the lone pair, while N-F bond are in opposite direction so they partly cuts the bond moment of lone pair.

Thus, dipole moment of NH_3 is higher than that of NF_3.

Q-24) What is meant by hybridisation of atomic orbitals? Describe the shapes of sp, sp^2, sp^3 hybrid orbitals.

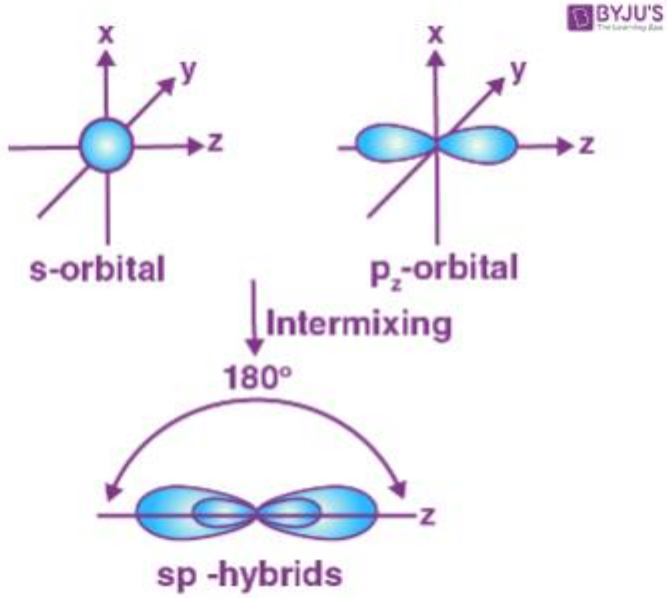
Ans.)

"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".

E.g. 1 s- orbital hybridises with 3 p- orbitals to form 4 sp^3 hybrid orbitals.

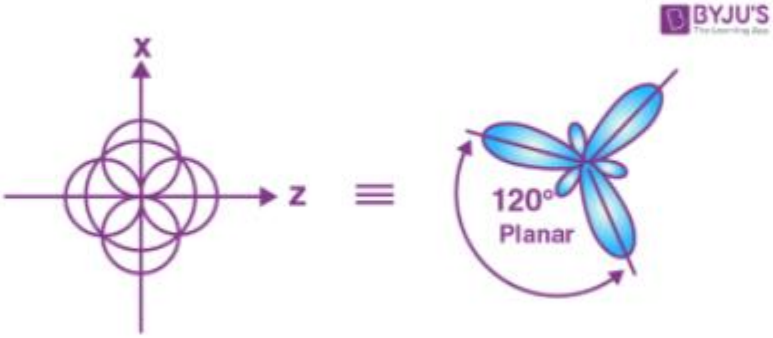
(a) sp hybrid orbital

1 s- orbital hybridises with 1 p- orbitals to form 2 sp hybrid orbitals. sp hybrid orbital is having linear shape. The formation of sp orbital is:



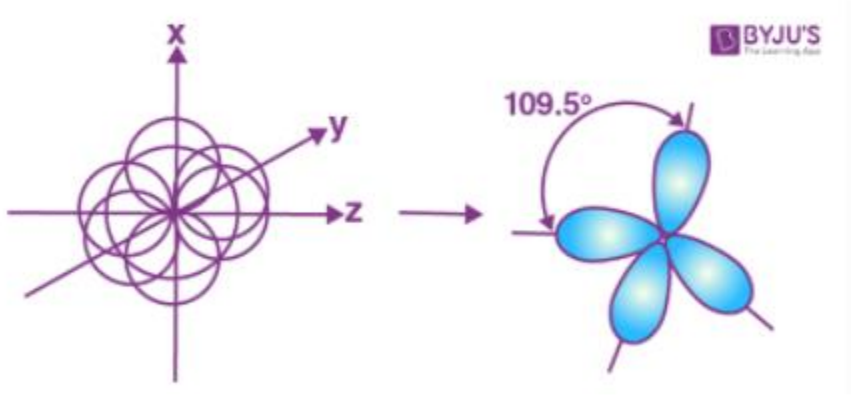
(b) sp^2 hybrid orbital

1 s- orbital hybridises with 2 p- orbitals to form 3 sp^2 hybrid orbitals. The shape of sp^2 orbital is trigonal planar.



(c) sp^3 hybrid orbital

1 s- orbital hybridises with 3 p- orbitals to form 4 sp^3 hybrid orbitals. The shape of sp^3 orbital is tetrahedron.

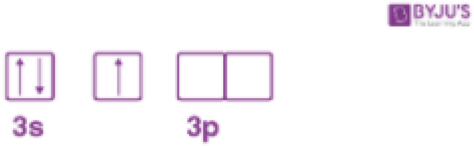


Q-25) Describe the change in hybridisation (if any) of the Al atom in the following reaction.

$AlCl_3$ + Cl^- –> $AlCl_4^-$

Ans.)

The ground state of valence orbital of Al –atom is:



In excited state the orbital picture of Al- atom is:

Thus, Al -atom in $AlCl_3$ undergoes sp^2 hybridisation and forms trigonal planar geometry. For the formation of $AlCl_4^-$ the vacant $3p_z$ orbital will also get involved. Thus, sp^2 hybridisation is converted into sp^3 hybridisation and forms a tetrahedral structure.

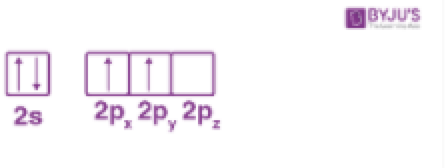
Q-26) Is there any change in the hybridisation of B and N atoms as a result of the following reaction?

$BF_3 + NH_3 \longrightarrow F_3B.NH_3$

Ans.)

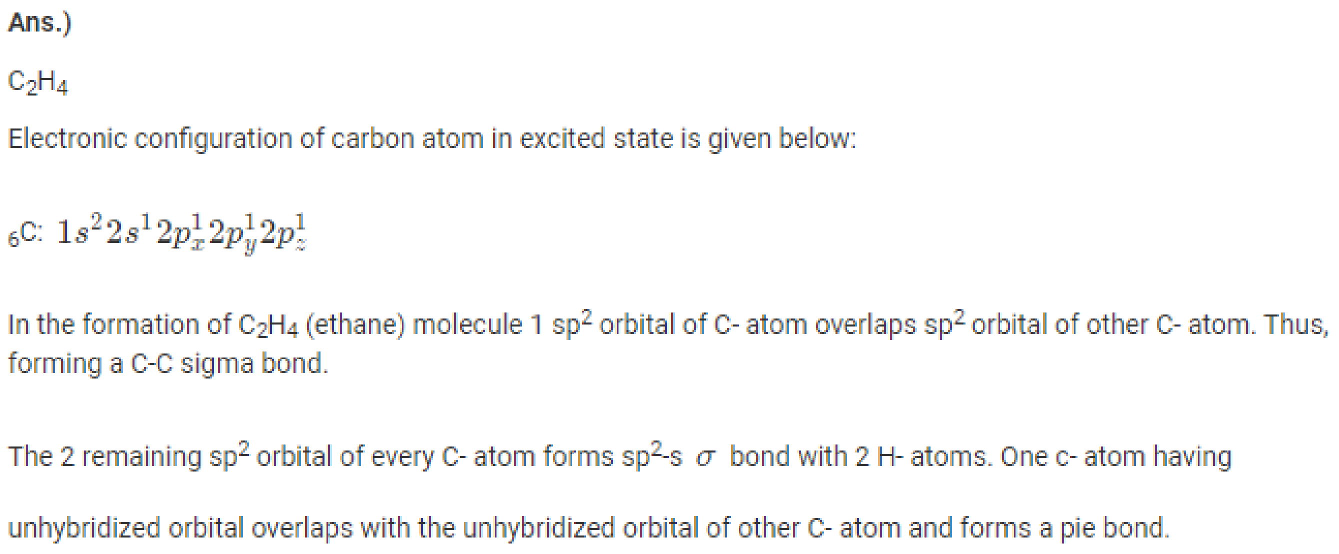
N- atom in NH_3 is having sp3 hybridization. Orbital picture of N- atom is shown below:

B- atom in NF_3 is having sp2 hybridisation. Orbital picture of B- atom is shown below:



On the reaction of NH_3 and BF_3, $F_3B.NH_3$ is obtained as product, as hybridization of B-atom is changed to sp^3. Although, hybridization of N- atoms remain unchanged.

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



Ans.)

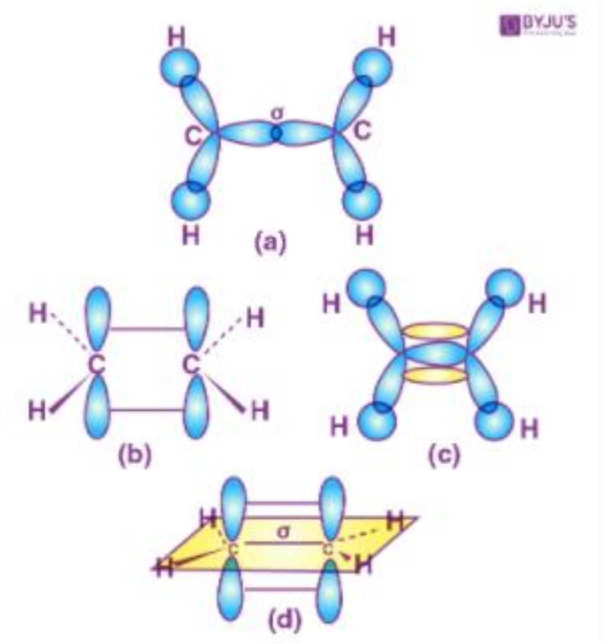
C_2H_4

Electronic configuration of carbon atom in excited state is given below:

$_6$C: $1s^2 2s^1 2p_x^1 2p_y^1 2p_z^1$

In the formation of C_2H_4 (ethane) molecule 1 sp^2 orbital of C- atom overlaps sp^2 orbital of other C- atom. Thus, forming a C-C sigma bond.

The 2 remaining sp^2 orbital of every C- atom forms sp^2-s σ bond with 2 H- atoms. One c- atom having unhybridized orbital overlaps with the unhybridized orbital of other C- atom and forms a pie bond.

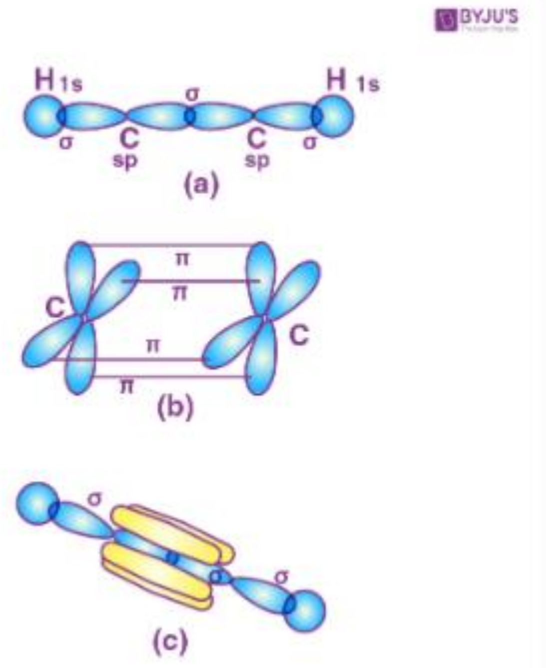


C_2H_2

In formation of ethyne(C_2H_2) molecule, C- atom is having sp hybridization with 2 2p- orbitals in unhybridized state.

1 sp orbital of each C- atom overlaps the inter-nuclear axis and forms C-C sigma bond. The 2nd sp orbital of each C- atom overlaps half-filled 1s orbital so as to form a sigma bond.

The triple bond between the 2 C- atoms has 1 sigma bond and 2 Pie bonds. This is because 2 unhybridized 2p-orbitals overlaps with the 2p- orbital of other C- atom, thus forming 2 pie bonds.

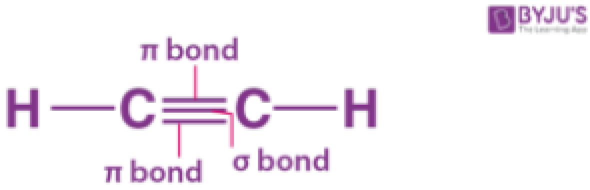


Q-28) What is the total number of sigma and pi bonds in the following molecules?
(a) C_2H_2
(b) C_2H_4

Ans.)

Single bond is formed as the axis of bonding orbital overlaps. Thus, it forms a σ bond. By sidewise overlapping of orbital double and triple bonds i.e. multiple bonds are formed. π bond is always present in the multiple bonds. Triple bond consist of 2 π and 1 σ bond.

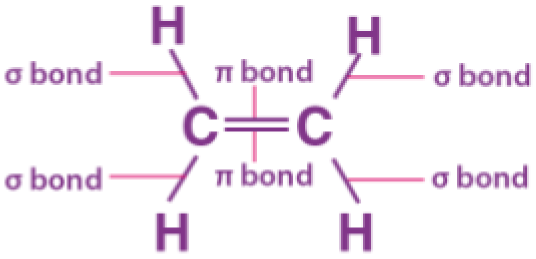
(a) C_2H_2



Thus, there are 2 π bonds and 3 σ in C_2H_2.

(b) C_2H_4

Thus, there are 1 π bonds and 5 σ in C_2H_4.



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

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

Ans.)

(c) $2p_y$ and $2p_y$

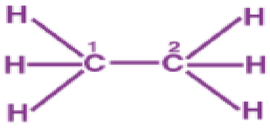
$2p_y$ and $2p_y$ orbitals won't form a σ as it will undergo lateral over lapping and will form a π bond.

Q-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_3CH_2-OH; (d) CH_3-CHO; (e) CH_3COOH.

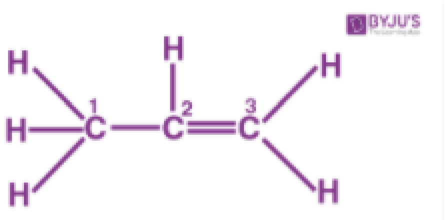
Ans.)

(a) CH_3-CH_3



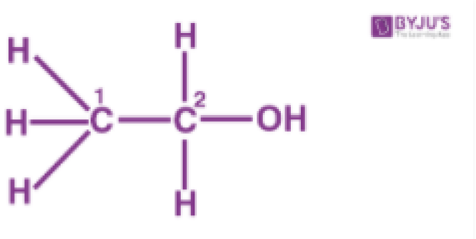
Here, C_1 and C_2 are having sp^3 hybridization.

(b) CH_3-CH=CH_2



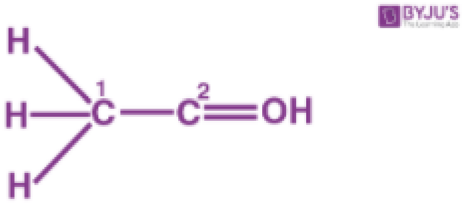
Here, C_3 and C_2 are having sp^2 hybridization and C_1 is having sp^3 hybridization.

(c) CH_3-CH_2-OH



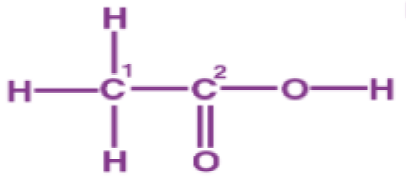
Here, C_1 and C_2 are having sp^3 hybridization.

(d) CH_3-CHO



Here, C_1 is having sp^3 hybridization and C_2 is having sp^2 hybridization.

(e) CH_3COOH



Here, C_1 is having sp^3 hybridization and C_2 is having sp^2 hybridization.

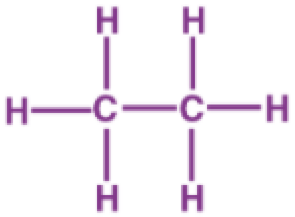
Q-31) What do you understand by bond pairs and lone pairs of electrons? Illustrate by giving one example of each type.

Ans.)

A covalent bond is formed when 2 atoms combine with each other by sharing their valence electrons.

“The shared pairs of electrons present between the bonded atoms are called bond pairs”. Each and every electron cannot participate in bonding. “The pairs of electrons which do not participate in bonding are called lone pairs”.

E.g. a) Ethane is having 7 bond pairs but zero lone pair.



b) Water is having 2 bond pairs and 2 lone pairs on O- atom.

Q-32) Distinguish between a sigma and a pi bond.

Ans.)

Sr. No.	Pi bond	Sigma bond
1	Pi bond is formed by lateral overlapping of orbitals.	Sigma bond is formed by end to end overlapping of orbitals.
2	It is comparatively weak bond.	It is comparatively strong bond.

3	There is only one overlapping orbital is p-p.	The overlapping orbitals are s-s, s-p, p-p.
4	Rotation around pi- bond is restricted.	Rotation is possible around sigma bond.
5	Electron cloud is not symmetrical about the line joining 2 nuclei.	Electron cloud is symmetrical about the line joining 2 nuclei.
6	It is having 2 electron clouds one above the plane of atomic nuclei and one below the plane of atomic nuclei.	It is having 1 electron cloud and that is symmetrical about the inter-nuclear axis.

Q-33) Explain the formation of H_2 molecule on the basis of valence bond theory.

Ans.)

Assuming 2 H- atoms X and Y with nuclei N_X and N_Y and electrons e_X and e_Y, respectively.

When X and Y are far for each other then there is no interaction between them. As soon as they come closer, the attractive force and repulsive force becomes active.

The repulsive forces are:

(i) Between electrons of both the atoms i.e. e_X and e_Y.

(ii) Between nuclei of both the atoms i.e. N_X and N_Y.

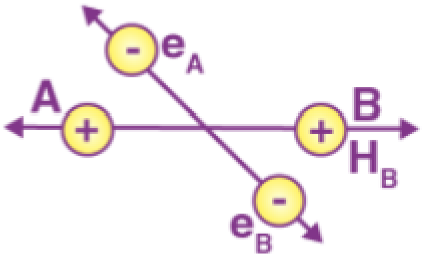
The attractive forces are:

(i) Between the electron and nucleus of the same atom i.e. N_X – e_X and N_Y -e_Y.

(ii) Between the electron of one atom and nucleus of other atom i.e. N_X–e_Y and N_Y–e_X.

The repulsive force pushes the 2 atoms apart whereas the attractive force tend to bring them together.

Repulsive forces:



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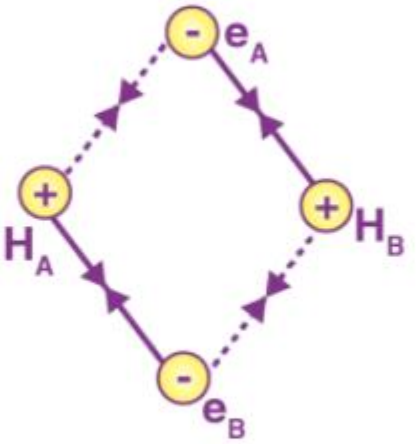
Attractive forces:



The values of repulsive forces are less than that of attractive forces. Thus, 2 atoms approach each other. Thus, there is a decrease in potential energy. At the end a stage is reached when the repulsive forces balance the attractive forces and the system achieves the minimum energy, which leads to formation of H2 molecule.

Q-34) Write the important conditions required for the linear combination of atomic orbitals to form molecular orbitals.

Ans.)

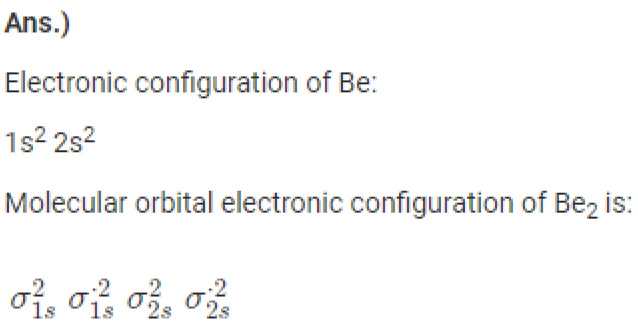
The condition that is required for linear combination of atomic orbitals to form molecular orbitals are as follows:

(i) The joining of atomic orbitals must have approximately the same energy. This implies in a homo-nuclear molecule, the 1s-orbital of one atom can join with the 1s- orbital of another atom, but cannot join with the 2s-orbital.

(ii) The joining atomic orbitals must have legitimate orientations to ensure the maximum overlap.

(iii) The overlapping must be to a large extent.

Q-35) Use molecular orbital theory to explain why the Be_2 molecule does not exist.



Ans.)

Electronic configuration of Be:

$1s^2 2s^2$

Molecular orbital electronic configuration of Be_2 is:

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

Thus, bond order of Be_2: $0.5(N_b - N_a)$.

N_b: No. of electrons in the bonding orbitals

N_a: No. of electrons in the anti-bonding orbitals

Therefore, bond order of $Be_2 = 0.5(4 - 4) = 0$

Zero value of bond order indicates that given molecule is unstable. Thus, Be_2 doesn't exist.

Q-36) Compare the relative stability of the following species and indicate their magnetic properties:

O_2, O_2^+ , O_2^- (Superoxide), O_2^{2-} (Peroxide)

Ans.)

O_2 contain 16 electrons i.e. 8 electrons from each O- atom.

- Electronic configuration of O_2 is:

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

As 1s- orbital of each O- atom does not involve in the bonding,

No. of bonding electrons = N_b = 8

No. of anti-bonding electrons = Na = 4

Now,

Bond order = 0.5(8 – 4) = 2

- Electronic configuration of O_2^+ is:

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

No. of bonding electrons = N_b = 8

No. of anti-bonding electrons = Na = 3

Now,

Bond order = 0.5(8 – 3) = 2.5

As, bonddissociation energy $\propto$ bond order

Hence, higher the bond order, higher stability will be there.

Arrangement according to decreasing order of stability is given as:

$$O_2^+ > O_2 > O_2^- > O_2^{2-}$$

- Electronic configuration of O_2^- (Superoxide) is:

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

No. of bonding electrons = N_b = 8

No. of anti-bonding electrons = Na = 5

Now,

Bond order = 0.5(8 − 5) = 1.5

- Electronic configuration of O_2^{2-} (peroxide) is:

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

No. of bonding electrons = N_b = 8

No. of anti-bonding electrons = Na = 6

Now,

Bond order = 0.5(8 − 6) = 1

Q-37) Write the significance of a plus and a minus sign shown in representing the orbitals

Ans.)

Generally, molecular orbital is represented by the ‘wave function’.

Positive (+) sign in representing a molecular orbital indicates positive wave function.

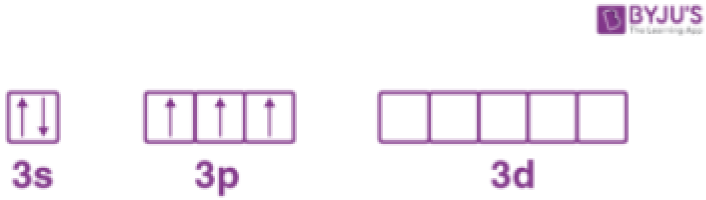
Negative (-) sign in representing a molecular orbital indicates negative wave function.

Q-38) Describe the hybridisation in case of PCl_5. Why are the axial bonds longer as compared to equatorial bonds?

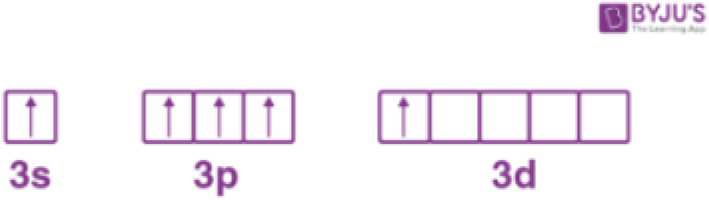
Ans.)

The electronic configuration of outer orbital of phosphorus in excited state and in ground state is given below:

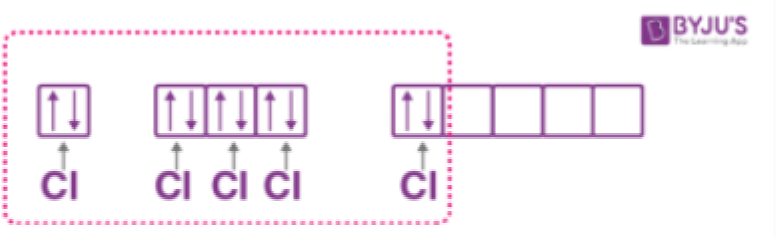
Ground State:



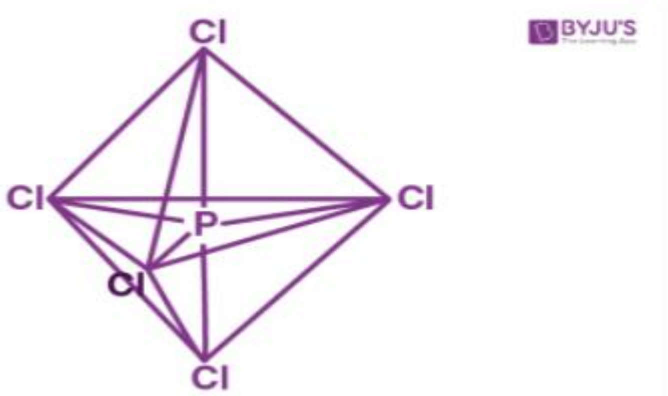
Excited State:



Phosphorus atom is having sp^3d hybridization. These orbitals are filled due to donation of electron pairs by 5 Cl- atoms: PCl_5



The 5 sp^3d hybrid orbitals present here are directed towards 5 corners of trigonal bipyramidal. Thus, the geometry of PCl_5 is given below:



PCl_5 contains 5 P- Cl sigma bonds. Out of which 3 P-Cl bonds lie in only 1 plane and they are making 120^0 with each other. And as these bonds lie in 1 plane they are known as equatorial bonds.

Out of 2 remaining P-Cl bonds, one bond lie above the equatorial plane and one bond lie below the equatorial bond. And they are making 90^0 with each other. These bonds are called axial bond.

Equatorial bond pairs repel axial bond pairs to a large extent so, equatorial bonds are slightly shorter than axial bonds.

Q-39) Define hydrogen bond. Is it weaker or stronger than the van der Waals forces?

Ans.)

"H-bond is defined as an attractive force acting between the hydrogen attached to an electronegative atom of one molecule and an electronegative atom of a different molecule".

As there is difference in the electronegativities between the atoms, thus, the bond pair electronegative atom and hydrogen atom is drifted away from H- atom. Therefore, hydrogen atom gets electropositive w.r.t. the other atom and procures a positive charge.

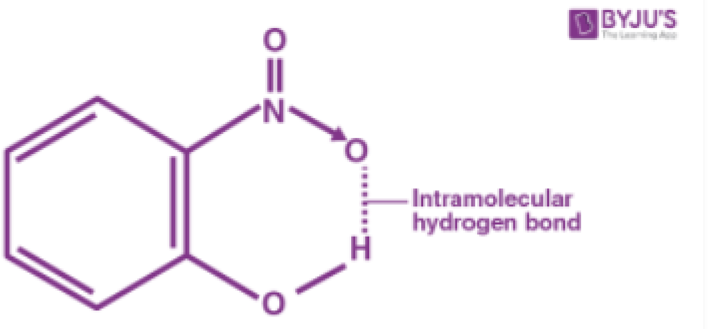
$$4^{\delta}-X^{\delta}\ \ldots\ldots\ldots\ldots H^{\delta+}-X^{\delta-}\ \ldots\ldots\ldots\ldots\ldots H^{\delta+}-X^{\delta-}$$

The value of H- bond is minimum in gaseous state and maximum in the solid-state.

Two types of hydrogen bonds are there:

(a) Intramolecular hydrogen bond e.g., o- nitrophenol

(b) Intermolecular hydrogen bond e.g., HF, H_2O etc.



H- bonds are stronger than Van der Waals forces as H- bond are regarded as extreme form of the dipole-dipole interaction.

Q-40) What is meant by the term bond order? Calculate the bond order of: N_2, O_2, O_2^+ and O_2^-.

Ans.)

Bond Order: It is defined as 0.5 times the difference between the "No. of electrons present in bonding orbitals and No. of electrons present in anti-bonding orbitals" of a molecule.

Bond Order = 0.5(N_b – N_a);

Na: No. of anti-bonding electrons

Nb: No. of bonding electrons

O_2 contain 16 electrons i.e. 8 electrons from each O- atom.

- Electronic configuration of O_2 is:

$[\sigma-(1s)]^2[\sigma^*(1s)]^2[\sigma(2s)]^2[\sigma^*(2s)]^2[\sigma(1p_z)]^2[\pi(2p_x)]^2[\pi(2p_y)]^2[\pi^*(2p_x)]^1[\pi^*(2p_y)]^1$ As 1s-orbital of each O- atom does not involve in the bonding,

No. of bonding electrons = N_b = 8

No. of anti-bonding electrons = Na = 4

Now,

Bond order = 0.5(8 – 4) = 2

- Electronic configuration of O_2^- (superoxide) is:

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

No. of bonding electrons = N_b = 8

No. of anti-bonding electrons = Na = 5

Now,

Bond order = 0.5(8 – 5) = 1.5

- Electronic configuration of O_2^+ is:

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

No. of bonding electrons = N_b = 8

No. of anti-bonding electrons = Na = 3

Now,

Bond order = 0.5(8 – 3) = 2.5

- Electronic configuration of N_2 is:

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

No. of bonding electrons = N_b = 10

No. of anti-bonding electrons = Na = 4

Now,

Bond order = 0.5(10 – 4) = 3