

→ The forces of attraction which holds the atoms together within a molecule

# CHEMICAL BONDING

\* How do atoms combine ?

## \* Important Terms

1. Valence electrons :- Number of electrons present in the valence shell.
  2. Bonding electrons :- Electrons taking part in the bond formation
  3. Non-Bonding electrons :- Electrons present in the valence shell not taking part in the bond formation [They are in pairs]
  4. Bond Pairs :- A pair of Two electrons involved in the covalent bond formation
  5. Lone Pairs :- A pair of Two electrons in the valence shell not involved in the bond formation.
- Most elements do not exist in the free atomic state (Eg:- H, O etc... occur as diatomic molecules)
  - Elements combine with each other to form Compounds. Such compounds occur as molecules.

## Chemical Inertness of Noble Gases

- Noble gases (He, Ne, Ar, Kr, Xe, Rn) do not form Compounds neither among themselves nor with other elements.
- Xenon however, forms fluoride and oxyfluoride Compounds under drastic conditions.

### Electronic Configuration of Noble gas elements.

	<u>Atomic Number</u>	<u>No. of valence electrons</u>
Helium	2	2
Neon	10	8
Argon	18	8
Krypton	36	8
Xenon	54	8
Radon	86	8

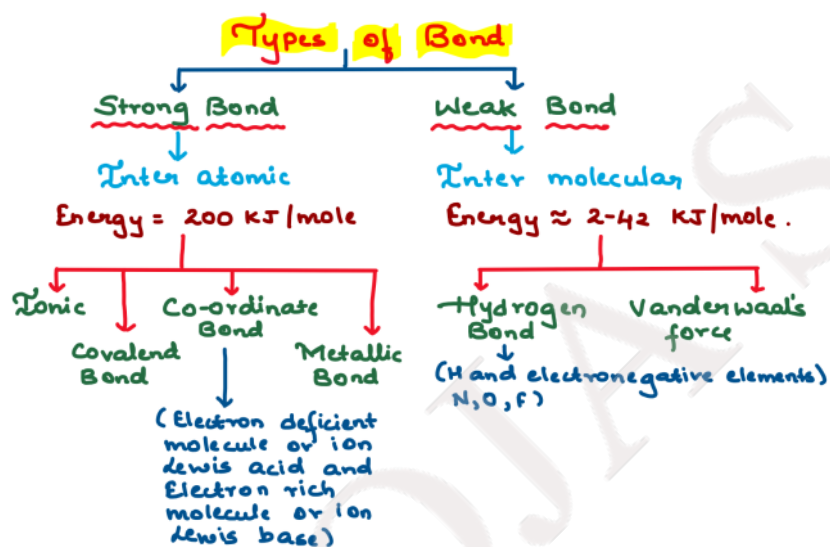
- From the e configuration given above, we see

- (i) Helium has only two electrons in its outermost shell.
- (ii) & all the other elements of this group (Ne, Ar, Kr, Xe, Rn) have eight electrons each in their outermost shell.

Thus,

- All noble gases have their outermost shell completely filled
- The atoms of all other elements which show chemical reactivity have less than eight electrons in their outermost shell.
- The atoms having a total of 2, 10, 18, 36, 54, 86 electrons are found to be most stable.
- Because of their stable electronic configuration, noble gases show no chemical reactivity
- The atoms of all other elements are not so stable and they tend to gain stability by acquiring an electronic configuration

- The atoms of all other elements are not so stable and they tend to gain stability by acquiring an electronic configuration of the nearest noble gas elements.
- A molecule will only be formed, if it is more stable and has lower energy than the individual atoms.
- The elements are divided into three classes:-
  - Electropositive elements:-** Which can lose one or more electrons easily.
  - Electronegative elements:-** Which can gain electrons. They have higher value of electronegativity.
  - Elements which have little tendency to lose or gain electrons

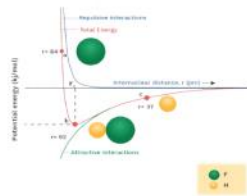


### Tendency to acquire minimum energy

- When two atoms comes closer, nucleus of one atom attracts the electron of another atom
- When net result is attraction, the total energy of the system (molecule) decreases and chemical bond forms.
- Bond formation is an **exothermic process**
  - Heat energy is given out.

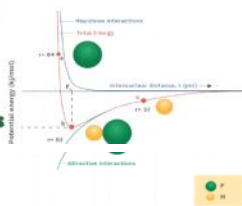
### How atom combine to form molecule.

(i) To obtain stable octet configuration or inert gas configuration



(ii) Decrease in potential energy - When two atoms approach to each other then there is force of repulsion due to nucleus of two atoms and also force of attraction between nucleus of one atom and electron of another atoms.

(iii) At point (b) force of attraction is dominating, so potential energy decreases and when potential energy becomes min, then bond is formed. When atom is made move closer, repulsion increases. Hence, after point (b) P.E increases.



(iv) During molecule formation attraction > repulsion. PE decreases.

(v) Between two atoms, when at any particular distance the PE is minimum or when they combine then the distance or length between them is termed as bond length.

\* The time at which molecule is formed there is decrease in potential energy.

## LEWIS REPRESENTATION

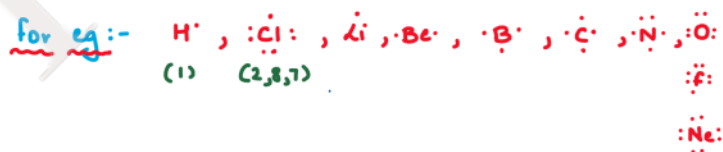
→ An American Chemist, G.N, Lewis introduced simple notation to denote the valence electrons in an atom.

→ These notations are called electron dot symbols or Lewis structures.

→ According to this method,

(i) The symbol of the element represents the nucleus alongwith all the inner electrons which do not take part in bond formation.

(ii) The dots on the symbol represent the valence electrons



(iii) The Lewis structure for chlorine molecule,

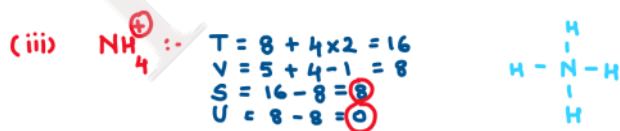
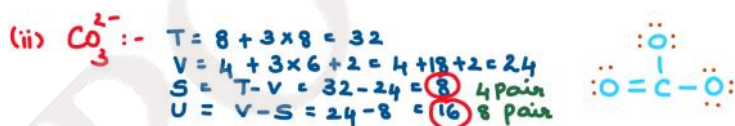
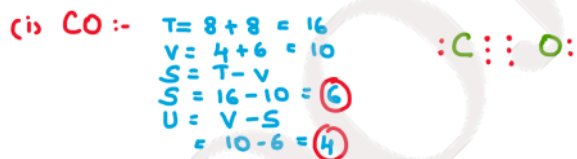




### STEPS OF WRITING LEWIS STRUCTURE FOR [V S O M L A C P]

- V (i) Count the number of valence electrons in each combining atoms in the molecule
- S (ii) Write skeletal structure of the molecule indicating valence electrons in each atom and formation of a single bond between the atoms.
- O (iii) If octet is not complete then form multiple bonds between the atoms so that octet of each bonding atom is complete.
- ML (iii) If anions, add extra electron for each negative charge.
- A (iv) If cations, subtract electron for each positive charge.
- C (v) In polyatomic molecules or ions, the electronegative atoms are surrounded around least electronegative atom.
- P (vi)

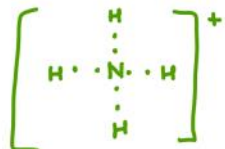
### Examples:-





\* Which atom in  $\text{NH}_4^+$  will have formal charge +1?

→ Lewis dot structure of  $\text{NH}_4^+$  is:-



formal charge on N atom =

$$\left[ \text{Total number of valence electrons in free N atom} \right] - \left[ \text{Total number of non-bonding electrons} \right]$$

$$- \frac{1}{2} \left[ \text{Total number of shared electrons in bonds} \right]$$

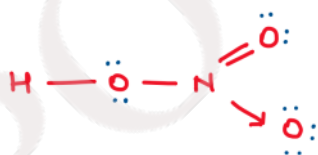
$$= 5 - 0 - \frac{1}{2}(8) = +1$$

Hence, N atom in  $\text{NH}_4^+$  has formal charge +1.

Q:- Calculate formal charge on

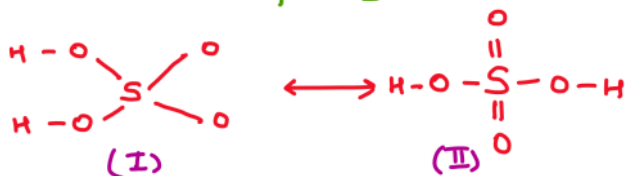
- (i) Nitrogen in  $\text{HNO}_3$
- (ii) Sulphur in  $\text{H}_2\text{SO}_4$
- (iii) Phosphorus in  $\text{PO}_4^{3-}$

(i) Lewis dot structure for  $\text{HNO}_3$  is,



$$\text{formal charge on 'N' atom} = 5 - 0 - \frac{1}{2} \times (8) = 1$$

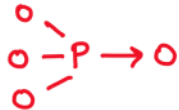
(2) Lewis dot structure for  $\text{H}_2\text{SO}_4$  is



Formal charge on 'S' atom =  $6 - 0 - \frac{1}{2} \times 8 = 2$   
(I)

(II) =  $6 - 0 - \frac{1}{2} \times 8$   
= 0

(3) Lewis dot structure for  $\text{PO}_4^{3-}$  is,



Formal charge on 'P' atom =  $5 - 0 - \frac{1}{2} \times 8 = +1$

The structure which has smallest formal charge has lowest energy and maximum stability

## Limitations of Octet Rule

Octet rule fails to explain the following:-

- 1) Incomplete octet :- The elements of Group 1, 2 and 13 contain less than 4 electrons in their valence shell. These elements cannot achieve an octet by electron sharing.  
(Electron deficient compounds)  
eg:-  $\text{BF}_3$ ,  $\text{BCl}_3$ ,  $\text{BeCl}_2$ .
- 2) Expanded octet :- The elements belonging to Group 15, 16 and 17 have more than 4 electrons in their outermost shell. The elements of these groups form stable compounds in which there are more than 8 electrons around the central atom.

eg:-  $\text{PF}_6$ ,  $\text{PCl}_5$

$\text{S}_2$ ,  $\text{SF}_6$

3) Odd electron molecule:- The molecules which do not obey rule and have odd number of electrons are called odd electron molecules.



## COVALENT

## BOND

A covalent bond is formed between two atoms (similar or dissimilar) by a mutual sharing of electrons.

A covalent bond is defined as the force of attraction arising due to mutual sharing of electrons between two atoms.

### Covalency

The number of electrons which an atom contributes towards mutual sharing during the formation of a chemical bond is called its covalency of in that compound.

Thus the covalency of hydrogen in H<sub>2</sub> → 1 (H-H)  
Oxygen in O<sub>2</sub> → 2 (O=O)  
Nitrogen in N<sub>2</sub> → 3 (N≡N)

### Comparison Between Single, Double, triple covalent Bonds

**Single Bond**:- It is formed by sharing of one electron pair (2e<sup>-</sup>)  
2e<sup>-</sup> hold the nuclei with a lesser force.  
Therefore, Bond length is maximum

**Double Bond**:- It is formed by sharing of Two electron pair (4e<sup>-</sup>)  
4e<sup>-</sup> hold the nuclei with little more force.  
Therefore, Bond length is lesser.

**Triple Bond**:- It is formed by sharing of Three electron pair (6e<sup>-</sup>)  
6e<sup>-</sup> hold the nuclei with greater force  
This decrease the distance of separation between two nuclei



Bond length is minimum

### Bond Length order

Triple Bond < Double Bond < Single Bond

### Bond energy order

Triple Bond > Double Bond > Single Bond

## VALENCE BOND THEORY

### \* Postulates:-

Valence Bond theory was proposed by It was further extended by Pauling 1931

Heitler in 1 and Slater

1926 in

- (1) A covalent bond is formed only when half filled orbit of two atoms overlap each other. In this process, energy of the system decreases.
- (2) Each overlapping atomic orbitals should contain an unpaired electron, with opposite spin.
- (3) On overlapping of orbitals, electron spins are neutral and electrons get paired.
- (4) The overlapping atomic orbitals must have nearly same energies.

orbitals

used

During the overlap of atomic orbitals, electron density increases between two nuclei and therefore repulsion between the two nuclei of bonded atoms decreases. This results in liberation of energy and increase in the attractive forces between the atoms.

(5)

A bond is formed at an equilibrium distance when the system has minimum potential energy and maximum stability.

(6)

The strength of a covalent bond depends upon the extent of overlap between the two atoms.

Greater the extent of overlap, more is the energy released and stronger is the bond formed.

(7)

Extent of overlapping depends on two factors.

Nature of orbitals :- p, d, f → directional orbitals

∴ s → non-directional  
↓  
more overlapping  
↓  
less overlapping

(i)

Nature of overlapping :-

co-axial overlapping → extent of overlapping is more  
collateral overlapping → extent of overlapping is less.

(ii)

\* Order of strength of co-axial overlapping



\* As the value of

'n' increases,

bond strength decreases.



\* If 'n' is same then,



(8) The number of covalent bonds formed by an atom element is equal to the number of unpaired electron present in the valence shell of the atom.

(9) Covalent bond formed by overlap of atomic orbital has directional characteristics as each atomic orbital except s-orbital has a particular direction in space.

(10) Geometry of a molecule is decided by direction, orientation of overlapping atomic orbitals.

1 of

is

,

2)

## Types of Overlapping

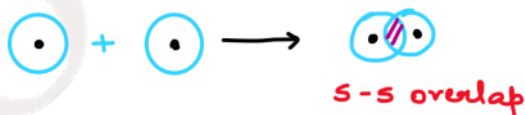
s-s overlap

s-p overlap

p-p overlap

In this type of overlap half-filled s-orbitals of the two combining atoms overlap each other.

(1) s-s overlap :-



In this type half-filled s orbital of one atom overlap with one of the half-filled p-orbital.



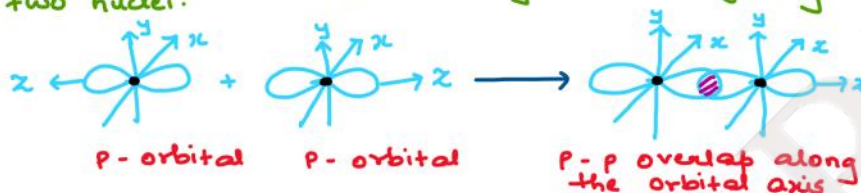
(2)

s-p overlap :-



### P-P overlap (Along the orbital axis) :-

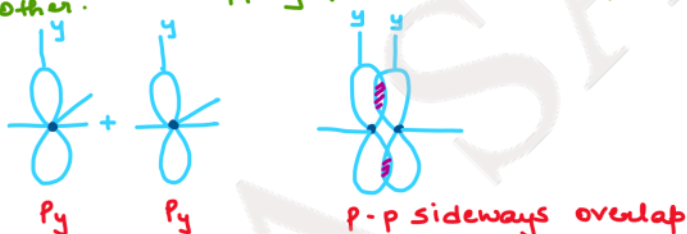
This is called head on, end on or end to end linear overlap. Here the overlap of the two half-filled p-orbitals takes place along the line joining the two nuclei.



(3)

### p-p Sideways overlap :- This is also called lateral overlap.

In this type of overlap, two p-orbitals overlap each other perpendicular to the internuclear axis. i.e. the two overlapping p-orbitals are parallel to each other.



(4)

### Types of covalent bonds

Sigma ( $\sigma$ )

Pi ( $\pi$ )

(1) Sigma ( $\sigma$ ) Bond :- Bond formed between the two atoms by the overlapping of half filled orbitals along their axis. (end to end overlap is called sigma bond.)

Eg:- The bond formed due to s-s, s-p, p-p overlap along the orbital axis.

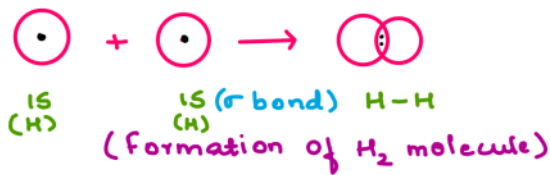
### \* Characteristics

- 1)  $\sigma$  bond is directional.
- 2)  $\sigma$  bond do not take part in resonance.
- 3) Free rotation is possible about a single  $\sigma$  bond.
- 4) Maximum overlapping is possible between electron clouds and hence it is a strong bond.
- 5) There can be only one  $\sigma$  bond between two atoms.

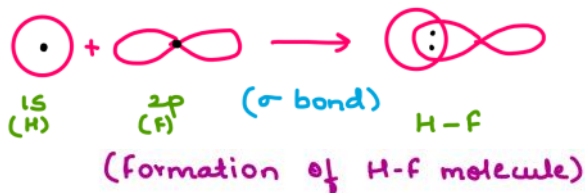
### overlapping :-

(1)

s-s



(2) S-p overlapping:-



(3) P-P overlapping (co-axial):-



The bond formed by sidewise (lateral) overlapping are known

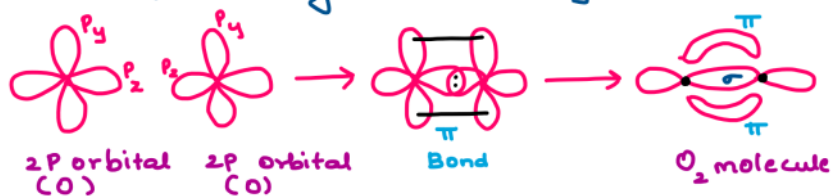
(2)



$\pi$  bonds

### \* Characteristics

- (1) Lateral overlapping is only partial, so formed are weaker and hence more reactive than  $\sigma$  bonds  
(Repulsion between the nucleus is more as orbitals have to come much close to each other for  $\pi$  bond formation)
- (2) Free rotation about a  $\pi$  bond is not possible.
- (3)  $\pi$  bond is weaker than  $\sigma$  bond
- (4)  $\pi$  bonds are less directional, so do not determine the shape of a molecule.
- (5)  $\pi$  bonds take part in resonance.
- (6)  $\pi$  bonds formed by pure or unhybrid orbitals.



# Hybridisation

\* Consider an example of Be compound:-

If it is formed without hybridisation then



Both the Be-Cl bonds should have different parameters and P-P bond strength > S-P bond strength.

But Practically bond strength and distance of both the Be-Cl bonds are same.

This problem may overcome if hybridisation of 's' and 'p' orbital occurs.

The concept of hybridisation is used to explain

- The nature of bonds
- Shape of the polyatomic molecules.

For an isolated atom, hybridisation has no meaning.

- According to the concept of hybridisation, certain atomic orbitals of nearly the same energy undergo mixing to produce equal number of new orbitals.
- The new orbitals so obtained are called hybrid orbitals.
- The process of mixing of the atomic orbitals to form new hybrid orbitals is called hybridisation.
- All hybrid orbitals of a particular kind have equal energy, identical shapes and are symmetrically oriented in space.

tals



Now after considering s-p hybridisation in  $\text{BeCl}_2$ , Bond strength of both the bonds will be equal.

## \* Characteristics

- (1) Hybridisation is a mixing of orbitals and not electrons. Therefore in hybridisation full filled, half filled and empty orbitals may take part.
- (2) Number of the hybrid orbitals formed is always equivalent to the number of atomic orbitals which have taken part in the process of hybridisation.
- (3) Each hybrid orbital having two lobes, one is larger and other is smaller. Bond will be formed from large lobe.



Structure of hybrid orbital.

The number of hybrid orbitals on central atom of a molecule or ion = number of  $\sigma$  bonds + lone pair of electron.

(4)

NOTE:- The 1<sup>st</sup> bond between two atoms will be sigma  
:- The other bond between same two atoms will be pi bond.  
:- The electron pair of an atom which do not take part in bond formation called as lone pair of electron.

(5) One element can represent many hybridisation state depending on experimental conditions  
for eg:-  $\rightarrow$  'C' showing  $sp$ ,  $sp^2$  and  $sp^3$  hybridisation in its compounds

(6) The repulsion between  $lp-lp > lp-bp > bp-bp$

(7) The directional properties in hybrid orbital is more than atomic orbitals. Therefore hybrid orbitals form stronger sigma bond.

The directional property of different hybrid orbitals will be in following order



### Steps Involved in hybridisation

(1) Formation of an excited state:- When the number of unpaired electrons present in the ground state of an atom of an element is not equal to (or less than) the valency of the element, an atom (in the ground state) absorbs energy and promotes one or more electrons from filled orbitals of lower energy to the vacant orbitals of higher energy, forming an excited state.  
 $\rightarrow$  In this state, the number of unpaired electrons is equal to the valency of the atom.

(2) Mixing and recasting of atomic orbitals:-

Depending upon the requirement of bonding, some or all the orbitals of the valence shell of the atom mix and reform a new set of equivalent orbitals having the same energy.

$\rightarrow$  These new orbitals formed are called hybridised orbitals.

$\rightarrow$  The number of hybridised orbitals is equal to the number of hybridising orbitals.

(3) Orientation of hybrid orbitals in space:- The hybrid orbitals orient themselves in space so as to minimise inter-electron repulsion to acquire stability. This arrangement gives a specific geometry to the molecule.

### Types of Hybridisation



#### (1) $sp$ Hybridisation

In this type of hybridisation, one 's' and one 'p' orbitals belonging to the same main energy level hybridise to give two  $sp$  hybrid orbitals.

## Types of Hybridisation

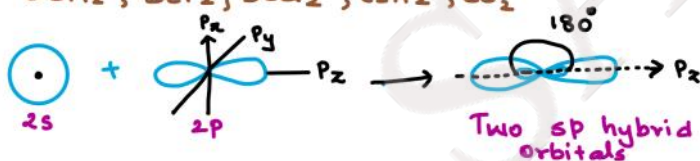


### (1) **Sp Hybridisation**

In this type of hybridisation, one 's' and one 'p' orbitals belonging to the same main energy level hybridise to give two sp hybrid orbitals.

- These sp hybrid orbitals are oriented at an angle of  $180^\circ$  to each other.
- Each hybrid orbital has 50% s and 50% p character.
- The other two p-orbitals remain unhybridised and are oriented at right angles to each other and to the internuclear axis.

Eg:-  $BeH_2$ ,  $BeF_2$ ,  $BeCl_2$ ,  $C_2H_2$ ,  $CO_2$



### **$BeF_2$ :-**

Beryllium difluoride has one beryllium atom and two fluorine atoms.

**Electronic Configuration:-**  $Be(4) :- 1s^2 2s^2$

$F(9) :- 1s^2 2s^2 2p_x^2 2p_y^2 2p_z^1$

The central beryllium atom undergoes sp-hybridisation.



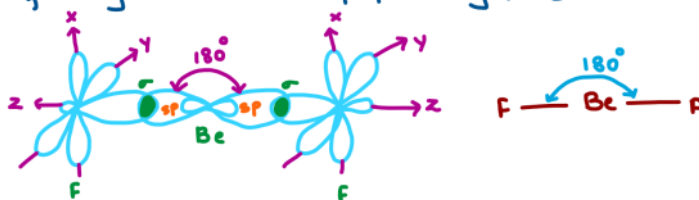
sp hybridised state

**Divalency of beryllium atom:-** The divalency of beryllium atom is explained by excitation of one of the paired electrons from lower energy 2s-orbital to the higher energy vacant  $2p_x$  orbital giving two unpaired electrons.

**sp hybridisation:-** In this, one s-orbital and one p-orbital of Be undergoes sp hybridisation forming two sp-hybridised orbitals which lie along the same axis on opposite sides with the angle  $180^\circ$ .



**formation of bonds:-** Two sp hybridised orbitals of beryllium atom overlap axially with half filled 2p orbitals of two fluorine atoms forming two Be-F (sp-p) sigma bonds.



**Geometry:-** Geometry of  $\text{BeF}_2$  molecule is linear or diagonal with F-Be-F bond angle  $180^\circ$ .

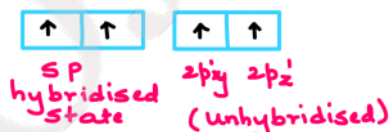
**Ethyne (Acetylene) ( $\text{C}_2\text{H}_2$ ):-**

Ethyne ( $\text{C}_2\text{H}_2$ ) molecule has two carbon atoms and two hydrogen atoms.

**Electronic Configuration:-**  
 $\text{H}(1):- 1s^1$   
 $\text{C}(6):- 1s^2 2s^2 2p_x 2p_y$

One of the paired electrons from 2s-orbital is excited to empty 2p-orbital giving four unpaired electrons for tetravalency of C-atom.

→ Each carbon atom undergoes sp-hybridisation



One s-orbital and one p-orbital of 'C' atom undergo hybridisation forming two sp-hybridised orbitals which lie along a straight line in opposite direction with angle  $180^\circ$ .

→ Two orbitals namely  $2p_y$  and  $2p_z$  of each carbon atom remain unhybridised.

**formation of bonds:-**

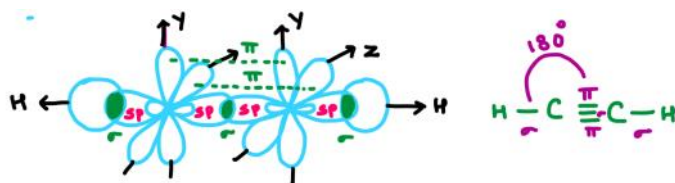
(a) One of the two sp-hybridised orbitals of one carbon atom overlaps axially with similar sp-hybrid orbital of another carbon atom to form C-C (sp-sp) sigma ( $\sigma$ ) bond.

(b) The remaining sp-hybridised orbital of one carbon atom overlaps axially with 1s-orbital of hydrogen atom forming two C-H (sp-s) sigma ( $\sigma$ ) covalent bonds.

(c) The unhybridised  $2p_y$  and  $2p_z$  orbitals of each carbon atom overlap laterally (parallelly) respectively to form two C-C pi ( $\pi$ ) bonds.

Thus in ethyne molecule, there are

- Two C-H sigma bonds
- One C-C sigma bond
- Two C-C pi bonds



**Geometry:-** The Geometry of Ethyne molecule is linear  
 → The bond angle is 180°.

**Q:-** In case of bond formation in acetylene molecule:-

- (a) How many covalent bonds are formed?
- (b) State no. of sigma and pi bonds formed
- (c) Name the type of hybridisation.

- Sol:-**
- (a) No. of covalent bonds = 5
  - (b) No. of sigma bonds = 3  
 No. of pi bonds = 2
  - (c) Type of hybridisation = sp

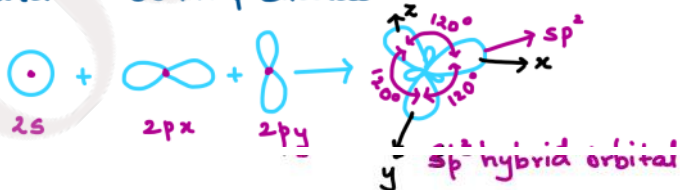
**sp<sup>2</sup>:-**

In this type of hybridisation, one s and two p orbitals of an atom undergo mixing to produce three equivalent sp<sup>2</sup> hybridised orbitals.

→ The three sp<sup>2</sup> hybrid orbitals are oriented in a plane along the three corners of an equilateral triangle i.e. they are at an angle of 120°.

→ The third p-orbital remains unchanged.

→ Each hybrid orbital has 33.3% s character and 66.7% p character.



Eg:- BF<sub>3</sub>, C<sub>2</sub>H<sub>4</sub>

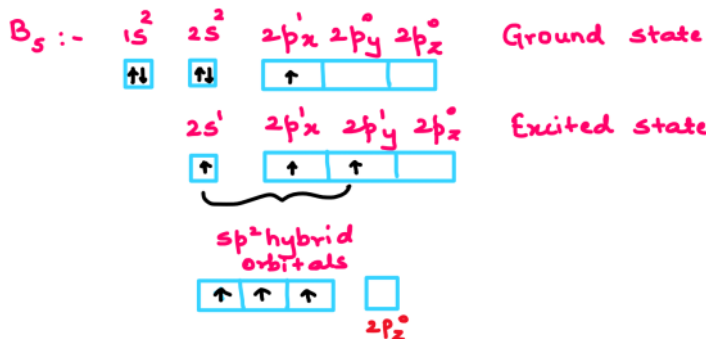
**BF<sub>3</sub>:-**

Boron trifluoride (BF<sub>3</sub>) molecule has one boron atom and three fluorine atoms.

**Electronic Configuration:-** F(9):- 1s<sup>2</sup> 2s<sup>2</sup> 2p<sup>2</sup> 2p<sup>2</sup> 2p<sup>1</sup>  
 B(5):- 1s<sup>2</sup> 2s<sup>2</sup> 2p<sup>1</sup>

→ One of the paired electrons from 2s-orbital is excited to empty 2p<sub>y</sub> orbital giving three unpaired electrons for trivalency of B.

→ The central boron atom undergoes sp<sup>2</sup> hybridisation.



**$sp^2$  hybridisation:** In this, one s-orbital and two p-orbitals of boron atom undergo  $sp^2$ -hybridisation forming three  $sp^2$  hybridised orbitals directed towards the three corners of an equilateral triangle with angle of  $120^\circ$ .



**Formation of bonds:** Each  $sp^2$  hybridised orbitals overlaps coaxially with  $2p_z$ -orbital of fluorine atom forming three  $sp^2$ -p sigma ( $\sigma$ ) bonds.

**Geometry:** Geometry of the molecule is trigonal planar.

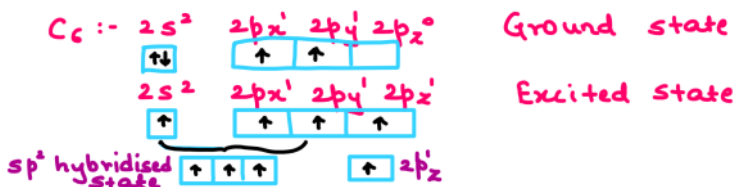
F-B-F bond angle is  $120^\circ$

### Ethene (Ethylene, $C_2H_4$ )

Ethene ( $C_2H_4$ ) (Ethylene) molecule has two carbon atoms and four hydrogen atoms.

**Electronic Configuration:**  $H_1$  :-  $1s^1$   
 $C_6$  :-  $1s^2 2s^2 2p_x^1 2p_y^1$

→ One of the Paired electrons from 2s-orbital is excited to the vacant  $2p_z$  orbital giving four unpaired electrons for tetravalency of C-atom. In this each carbon atom undergoes  $sp^2$  hybridisation.



**Need of hybridisation:** The observed H-C-H bond angle and formation of  $\pi$  bond between two carbon atoms can be explained by  $sp^2$  hybridisation.

**$sp^2$  hybridisation:** In this, one s-orbital and two p-orbitals of carbon atom undergo  $sp^2$  hybridisation forming three  $sp^2$  hybridised orbitals directed towards the three corners of an equilateral triangle with the angle of  $120^\circ$ .  $2p_z$ -orbital remains unhybridised.

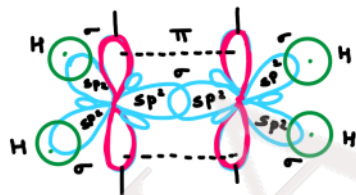
orbitals directed towards the three corners of an equilateral triangle with the angle of  $120^\circ$ .  $2p_z$ -orbital remains unhybridised.

### Formation of bonds:-

- Each  $sp^2$ -hybridised orbital of each carbon atom overlaps coaxially with  $1s^1$  orbitals of two hydrogen atoms forming two C-H  $sp^2-s$  sigma ( $\sigma$ ) bonds.
- The remaining  $sp^2$  hybridised orbital of one carbon atom overlaps axially with  $sp^2$  hybridised orbital of another carbon atom to form a C-C ( $sp^2-sp^2$ ) sigma ( $\sigma$ ) bond.
- The unhybridised half filled  $2p_z$ -orbitals of two carbon atoms, having electrons with opposite spins overlap laterally to form a C-C ( $2p_z-2p_z$ ) pi ( $\pi$ ) bond.

Thus in ethene molecule, there are

four C-H sigma bonds  
 one C-C sigma bond  
 one C-C pi bond.



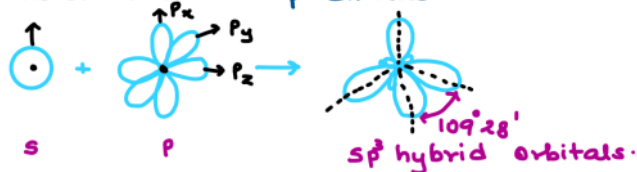
**Geometry:-** Geometry of ethene molecule is trigonal planar.

- All four C-H bonds are coplanar.
- The H-C-H bond angle is  $120^\circ$ .

### $sp^3$

In this type of hybridisation, one 's' and three 'p' orbitals of an atom undergo mixing to produce four equivalent  $sp^3$  hybridised orbitals.

- The four  $sp^3$  hybrid orbitals are directed along the four corners of a tetrahedron and are inclined to each other at an angle  $109^\circ 28'$ .
- Each of these hybrid orbitals has 25% s-character and 75% p-character.

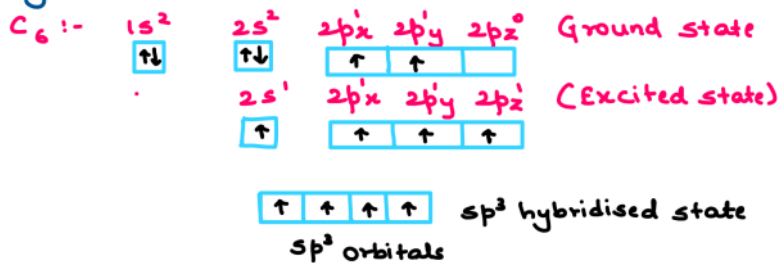


### Methane ( $CH_4$ )

Methane ( $CH_4$ ) molecule has one carbon atom and four hydrogen atoms.

**Electronic Configuration of H(1):-**  $1s^1$   
**C(6):-**  $1s^2 2s^2 2p_x^1 2p_y^1$

The central carbon atom undergoes  $sp^3$  hybridisation.



**$sp^3$ -hybridisation:** In this, one s-orbital and three p-orbitals of C-atom undergo  $sp^3$  hybridisation forming four equivalent  $sp^3$  hybridised orbitals directed towards four corners of tetrahedron with the angle  $109^\circ 28'$ .



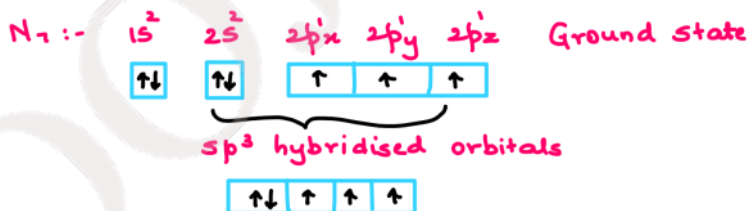
**Geometry:-** The Geometry of methane is Tetrahedral

**Ammonia ( $NH_3$ ):-**

Ammonia molecule has one nitrogen atom and three hydrogen atoms.

**Electronic Configuration of  $H(1)$ :-**  $1s^1$   
 **$N(7)$ :-**  $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$

The central nitrogen atom undergoes  $sp^3$  hybridisation.



**$sp^3$  hybridisation:** In this, one s-orbital and three p-orbitals of N atom undergo  $sp^3$  hybridisation forming four  $sp^3$  hybridised orbitals directed towards the four corners of regular tetrahedron.

→ One of the  $sp^3$ -hybrid orbitals has a lone pair of electrons.



**Geometry:-** In ammonia molecule, one of the  $sp^3$  orbitals has a lone pair of electrons, due to which, there is a repulsion between lone-pair and bonding pair of electrons.  
 → As a result, the H-N-H bond angle is reduced from regular tetrahedral angle  $109^\circ 28'$  (Lone Pair-bond Pair > bond pair-bond pair) to  $107^\circ 18'$ .

**Geometry:-** In ammonia molecule, one of the  $sp^3$  orbitals has a lone pair of electrons, due to which, there is a repulsion between lone-pair and bonding pair of electrons.

→ As a result, the H-N-H bond angle is reduced from regular tetrahedral angle  $109^\circ 28'$  (Lone Pair-bond Pair > bond pair-bond pair) to  $107^\circ 18'$ .

→ The Geometry of the ammonia molecule is distorted tetrahedral or Pyramidal.

**Water ( $H_2O$ ):-**

Water molecule has one oxygen atom and two hydrogen atoms.

**Electronic configuration** H(1):-  $1s^1$   
O(8):-  $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$

The central oxygen atom undergoes  $sp^3$ -hybridisation

O<sub>8</sub>:-  $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$  Ground state

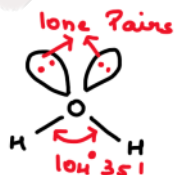


$sp^3$  hybridised state



**$sp^3$  hybridisation:-** In this, one s-orbital and three p-orbitals of Oxygen atom undergo  $sp^3$  hybridisation forming four  $sp^3$ -hybridised orbitals directed towards four corners of a tetrahedron.

→ Two of the hybrid orbitals have lone-pair electrons



**Geometry:-** In water molecule, two of the  $sp^3$  hybridised orbitals have lone pairs of electrons due to which there is force of repulsion between lone pairs and bonding pairs.

→ These forces cause distortion in the H-O-H bond angle which is reduced from the expected regular tetrahedral angle  $109^\circ 28'$  to  $104^\circ 35'$ .

→ Hence the geometry of water molecule is angular or 'V' shaped.

— x —