CHEMICAL BONDING

Noble Gases-Octet Rule

Noble gases are also known as inert gases or rare gases. These gases are stable and exist in free state because they have their outermost orbit completely filled. Hence, they are chemically unreactive or intert.

Electronic configurations of noble gases

Element	Symbol	Atomic number	Electronic configuration	Electronic structure
Helium	He	2	2	$1s^2$
Neon	Ne	10	2,8	$1s^2 2s^2 2p^6$
Argon	Ar	18	2, 8, 8	$1s^2 2s^2 2p^6 3s^2 3p^6$
Krypton	Kr	36	2, 8, 18, 8	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$
Xenon	Xe	54	2, 8, 18, 18, 8	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6$
Radon	Rn	86	2, 8, 18, 32, 18, 8	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 5d^{10} 6s^2 6p^6$

All noble gases (except Helium) have eight electrons in their outermost shell. This led to idea of octet rule which states that *atom tends to lose, gain or share valence electrons with other atom to achieve the state of noble gases.*

Causes of Chemical Combination

The cause of chemical combinatioin is to gain the state of noble gases and thus the state of lower energy, extra stability, low electron affinity, high ionisation energy and chemical inertness. This can be done:

(a) By loss or gain of valence electrons:
Sodium atom loses one elctron to attain the configuration of neon atom:

the configuration of neon atom:

Na	$\rightarrow$	Na^+	$+$	e^-
sodium atom (2, 8, 1)		Sodium ion (2, 8)		Electron

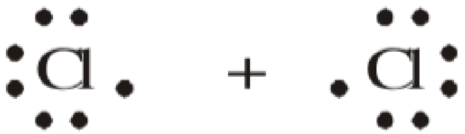
Chlorine atom gains one electron to attain the configuration of argon atom:

Cl	$+$	e^-	$\rightarrow$	Cl^-
Chlorine atom (2, 8, 7)		Electron		Chloride ion (2, 8, 8)

Hence,

Na^+	$+$	Cl^-	$\rightarrow$	$NaCl$
Sodium ion		Chloride ion		Sodium chloride

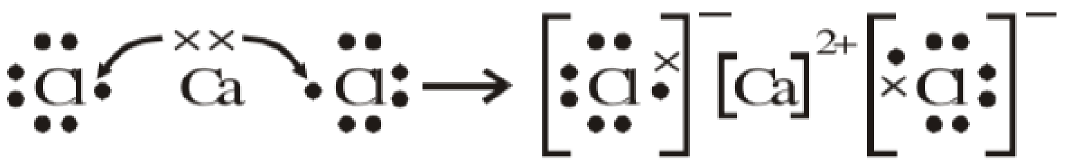
(b) By sharing of valence electrons with other atom:



Chlorine atom Chlorine atom Chlorine molecule

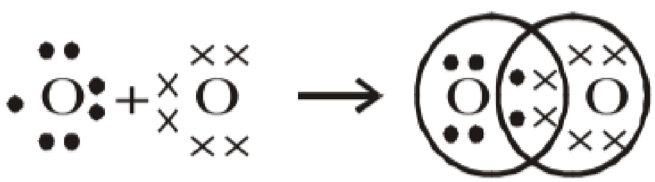
During the formation of chemical bond, energy is released and so the resultant molecule has lower energy and greater stability.

1. Ionic or electrovalent bonding:

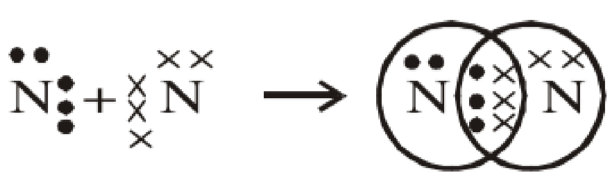


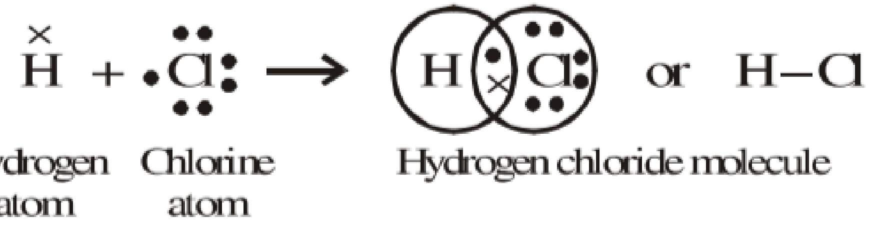
Chlorine atom (2, 8, 7)	Calcium atom (2, 8, 8, 2)	Chlorine atom (2, 8, 7)	Chloride ion (2, 8, 8)	Calcium ion (2, 8, 8)	Chloride ion (2, 8, 8)

2. Covalent bonding:

(*i*)  or $O=O$

Oxygen atom Oxygen atom Oxygen molecule

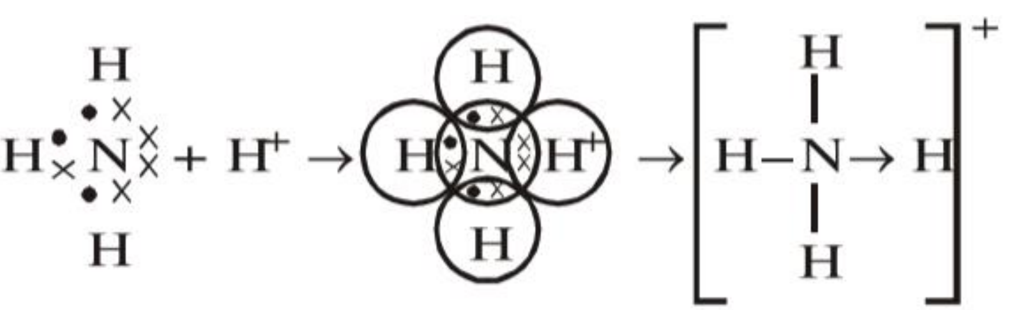
(*ii*)  or $N\equiv N$

Nitrogen atom Nitrogen atom Nitrogen molecule (*iii*) 

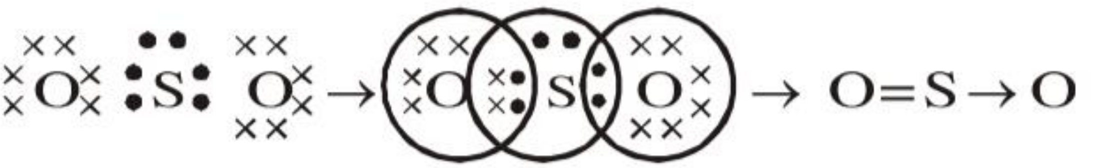
Co-ordinate Covalent Bond

In this bond shared pair of electrons (lone-pair) comes from one atom only. The atom which donates lone pair is known as donor and other atom which accepts, it is known as acceptor. It is represented by an arrow (→) from doner to acceptor atom.

(*a*) Formation of ammonium ion



(*b*) Formation of sulphur dioxide

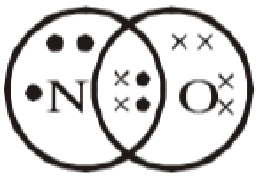


Deviation from Octet Rule

In some cases, it was observed that atoms do not follow octet rule during bonding. Instead of having followed octet rule, the molecules have,

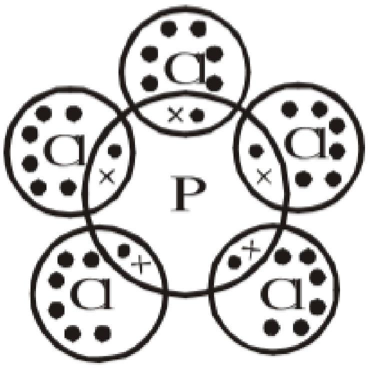
(*a*) odd number of electrons.
(*b*) expanded octet.
(*c*) incomplete octet.

(*a*) Odd number of electrons: Normally molecules have even number of electrons and thus complete pair of electrons. But some molecules like ClO_2, NO and NO_2 have odd number of electrons. Such as nitric oxide



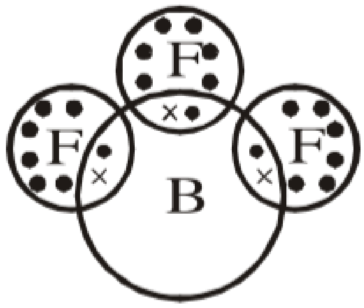
N has 7 electrons in outermost orbit.

(*b*) Expanded Octet: Some compounds have more than 8 electrons in outermost shell such as Phosphorus pentachloride.



P has 10 electrons in its valence shell.

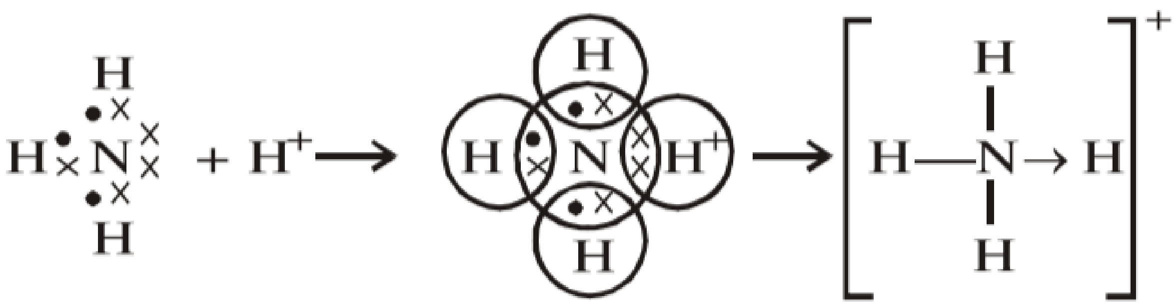
(*c*) Imcomplete octet: Some compounds are fewer number of electrons than 8. Such as Boron trifluoride.



BF_3 has 6 electrons in outermost shell of B atom.

3. Polar covalent bond:

4. Coordinate covalent bond:



5. Electropositive element + Electronegative element → Ionic bond

   Electronegative element + Electronegative element → Covalent bond

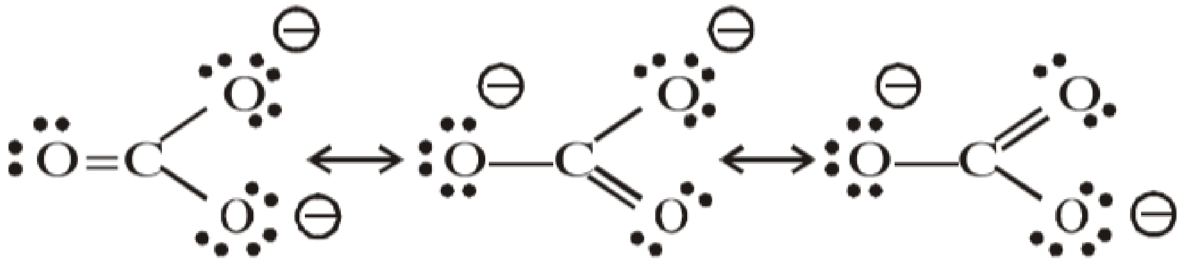
   Electropositive element + Electropositive element → Metallic bond

Resonance

Sometimes molecule or ion is represented by more than one electronic structure, in which only one represents all the properties of that molecule or ion.

Various dot structures for a molecule are known as resonating structures. Actual structure is intermediate of all these resonating structures *e.g.* :

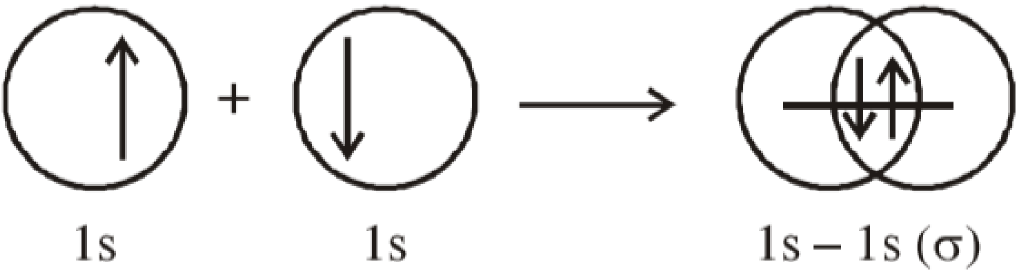
Resonance forms of carbonate ion, CO_3^{2-}



Sigma (σ) and pi (π) Bonding

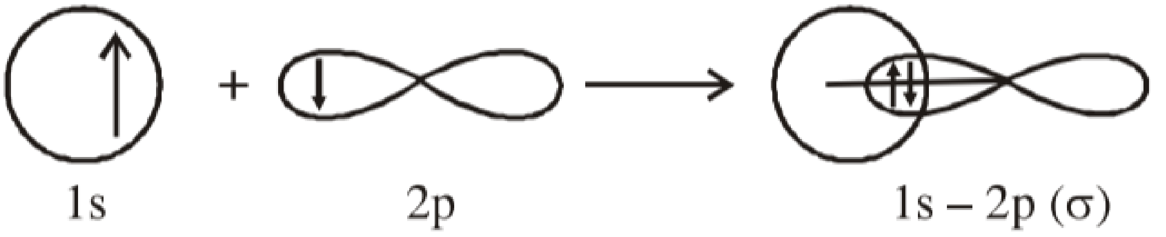
Sigma (σ) **bond** between two atoms involves head on overlapping along their internuclear axis which gives rise to maximum electron density on the axis. Overlapping takes place in the following ways:

(*i*) **s-s overlapping:** This type of overlapping takes place between half filled s-orbitals of the atoms, *e.g.* in H_2 molecule.



Formation of hydrogen molecule

(*ii*) **s-p overlapping:** This involves overlapping between half filled s-orbital of one atom with half filled p-orbital of another atom. *e.g.* In HF molecule.

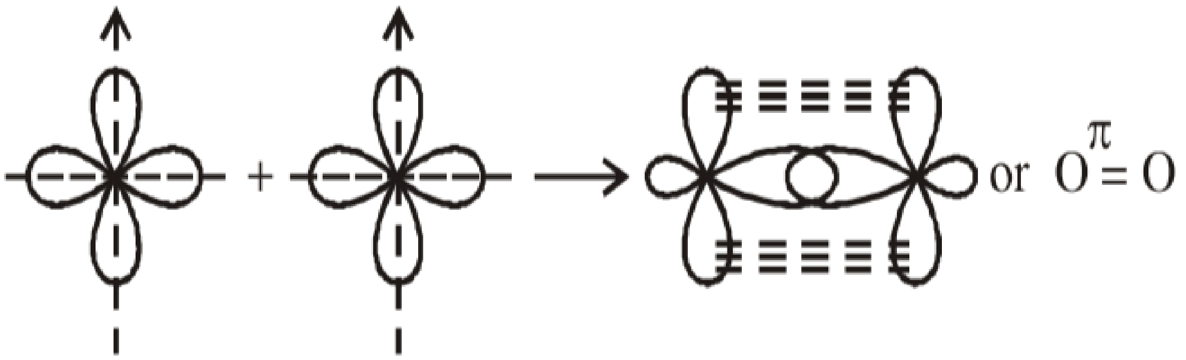


Formation of hydrogen fluoride molecule

(*iii*) **p-p overlapping:** This involves end to end overlapping of p-orbitals of different atoms. e.g. in F_2 molecule.

Formation of fluorine molecule

Pi (π) bonding: pi bond is formed by lateral or side ways overlapping of p-orbitals i.e. by overlapping of p-orbitals in a direction at right angles to the internuclear axis. e.g. O_2 molecule.



Two half filled p-orbitrals of O-atom *Oxygen molecule*

Types of Hybrid Orbitals

(*i*) **sp *hybridization*:** One s orbital and one p orbital hybridize forming linear shaped molecule of bond angle 180° such as BF_2.

(*ii*) **sp^2 hybridization:** One s orbital and two p orbitals hybridize to form molecule planar triangular in shape with bond angle of 120°. such as C_2H_2, N_2O, BCl_3, SO_3.

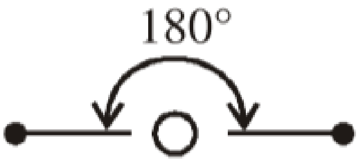
(*iii*) **sp^3 hybridization:** One s orbital and three p orbitals hybridize to form molecule having bond angle of 109° 28′ and having shape tetrahedral.

(*iv*) **sp^3 d hybridization:** One s, three p and one d orbitals hybridize to give sp^3d

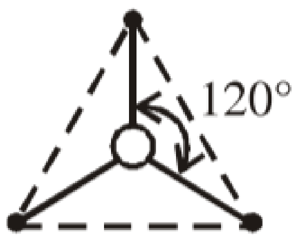
hybridization. It has trigonal bipyramidal shape. Such as PCl_5, SF_4, XeF_2, ICl_3, ClF_3.

(*v*) sp^3d^2 hybridization: One s, 3p and 2d orbitals are hybridized. Shape of the molecule is octahedral. It has 6 hybridized orbitals. e.g. SF_6.

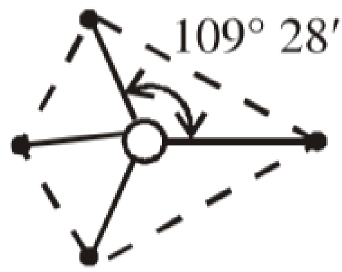
(*vi*) sp^3d^3 hybridization: One s, 3p and 3d orbitals are hybridized to give 7 hybridized orbitals. Five are coplanar and two are perpendicular to this plane. e.g. Cl_7I.



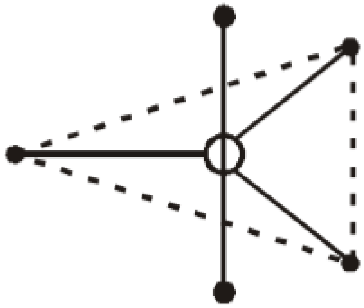
Linear shape
(sp molecule)

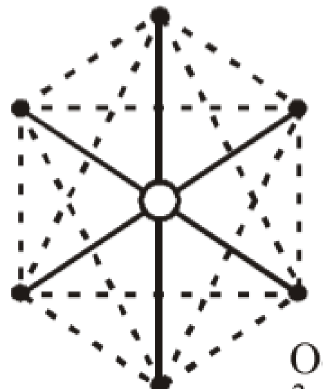


Trigonal planer
(sp^2 molecule)



Tetrahedral
(sp^3 molecule)


Trigonal bipyramidal
(sp^3d molecule)


Octahedral
(sp^3d^2 molecule)

Difference Between Sigma and pi-Bonds

Sigma (σ) Bond	Pi (π) Bond
I. It is formed by end to end overlapping of orbitals.	I. It is formed by side ways overlapping of orbitals.
II. σ-bond is stronger	II. π-bond is weaker.
III. This is formed by overlapping between s-s, s-p and p-p orbitals.	II. This is formed by overlapping between p-p orbitals only.
IV. It exists alone or along π-bond.	IV. It always exists along σ-bond.
V. Electron cloud is symmertical about the axis.	V. Electron cloud is unsymmertical.
VI. Free rotation about σ-bond is possible.	VI. Free rotation about π-bond is not possible, as it leads to breakage of bond.
VII. It consists of only one electron cloud about the intern-uclear axis.	VII. it consists of two electron clouds one above and other below the plane.

Some Common Bond Length

Bond	Bond Length (Å)
C–C (Alkane)	1.54
C=C (Alkene)	1.34
C≡C (Alkyne)	1.20
C–H (Alkane)	1.09
C=C (Benzene)	1.39
O–H (Alcohols)	0.96
C–O (Alcohols)	1.43
C=O (Ketones)	1.21
C–Cl (Chloroalkane)	1.77
N–H (Amines)	1.02
C–N (Amines)	1.47
C=N (Isocyanides)	1.52
C≡N (Cyanides)	1.15
C–Br	1.91
C–I	2.12
H–H	0.74
N≡N	1.09

Bond	Bond Length (Å)
O=O	1.21
F–F	1.42
Cl–Cl	1.99
Br–Br	2.28
I–I	2.67
H–F	0.92
H–Cl	1.27
H–Br	1.41
H–I	1.61