Government Arts College,CBE

Department of chemistry
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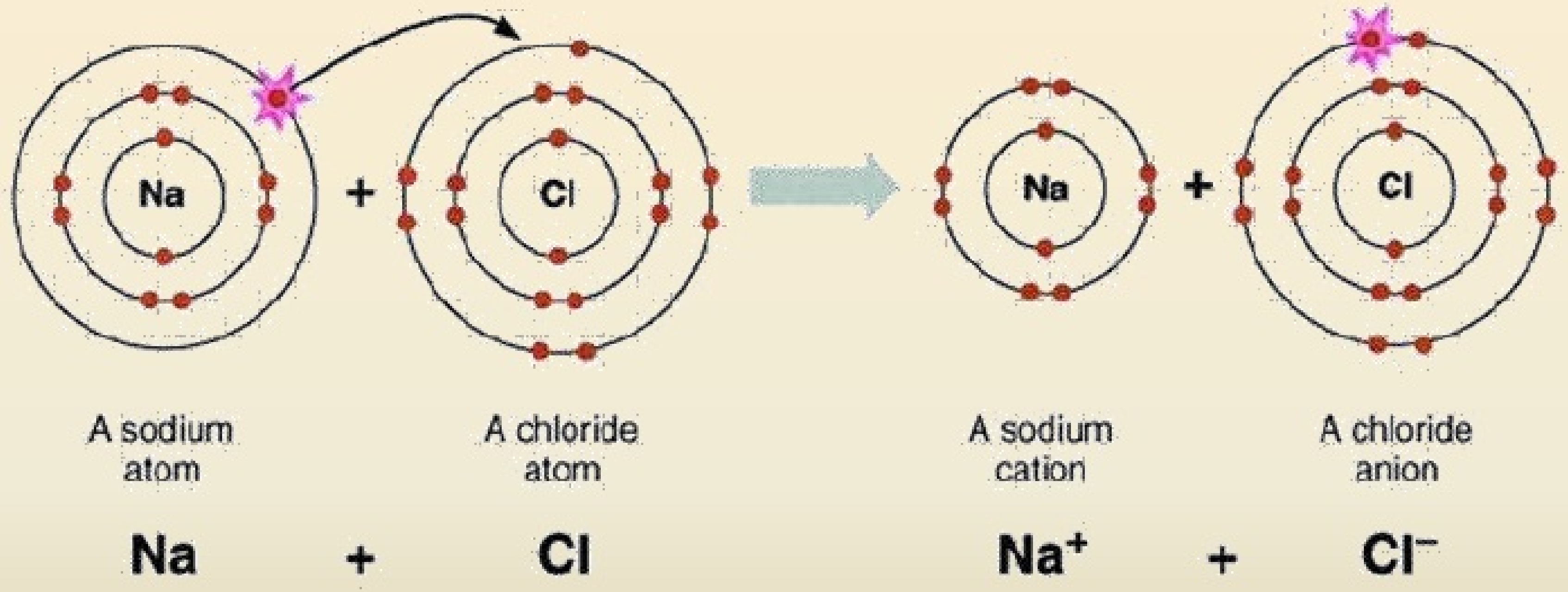
Unit-I
CHEMICAL BONDING

- A **chemical bond** - attraction between atoms, ions or molecules that enables the formation of **chemical** compounds.
- The **bond** may result from the electrostatic force of attraction between oppositely charged ions as in ionic **bonds**
- The sharing of electrons as in covalent **bonds**.

Ionic Bonding



CA Standards

- *Students know* atoms combine to form molecules by sharing electrons to form covalent or metallic bonds or by exchanging electrons to form ionic bonds.
- *Students know* salt crystals, such as NaCl, are repeating patterns of positive and negative ions held together by electrostatic attraction.

Bonds

- Forces that hold groups of atoms together and make them function as a unit.
  - Ionic bonds – transfer of electrons
  - Covalent bonds – sharing of electrons

The Octet Rule – Ionic Compounds

Ionic compounds form so that each atom, by *gaining or losing* electrons, has an octet of electrons in its highest occupied energy level.

Metals lose electrons to form positively-charged cations

Nonmetals gains electrons to form negatively-charged anions

Ionic Bonding:
The Formation of Sodium Chloride

- Sodium has 1 valence electron
- Chlorine has 7 valence electrons
- An electron transferred gives each an octet

Na: $1s^2 2s^2 2p^6 \mathbf{3s^1}$

Cl: $1s^2 2s^2 2p^6 3s^2 \mathbf{3p^5}$

Ionic Bonding: The Formation of Sodium Chloride

This transfer forms ions, each with an octet:

Na^+ $1s^2 2s^2 2p^6$

Cl^- $1s^2 2s^2 2p^6 3s^2 3p^6$

Ionic Bonding: The Formation of Sodium Chloride

The resulting ions come together due to electrostatic attraction (opposites attract):

Na^{+} Cl^{-}

The net charge on the compound must equal zero

Examples of Ionic compounds

$Mg^{2+}Cl^{-}_{2}$ **Magnesium chloride:** Magnesium loses two electrons and each chlorine gains one electron

$Na^{+}_{2}O^{2-}$ **Sodium oxide:** Each sodium loses one electron and the oxygen gains two electrons

$Al^{3+}_{2}S^{2-}_{3}$ **Aluminum sulfide:** Each aluminum loses two electrons (six total) and each sulfur gains two electrons (six total)

Metal	Monatomic Cations	Ion name
Lithium	Li^{+}	Lithium
Sodium	Na^{+}	Sodium
Potassium	K^{+}	Potassium
Magnesium	Mg^{2+}	Magnesium
Calcium	Ca^{2+}	Calcium
Barium	Ba^{2+}	Barium
Aluminum	Al^{3+}	Aluminum

Recall……..

IONIC BOND- Transfer of electrons

Ex- NaCl

COVALENT BOND- Sharing of electrons,

Ex-Cl-Cl

METALLIC BOND –Exchangeable of electrons

Properties of Ionic Compounds

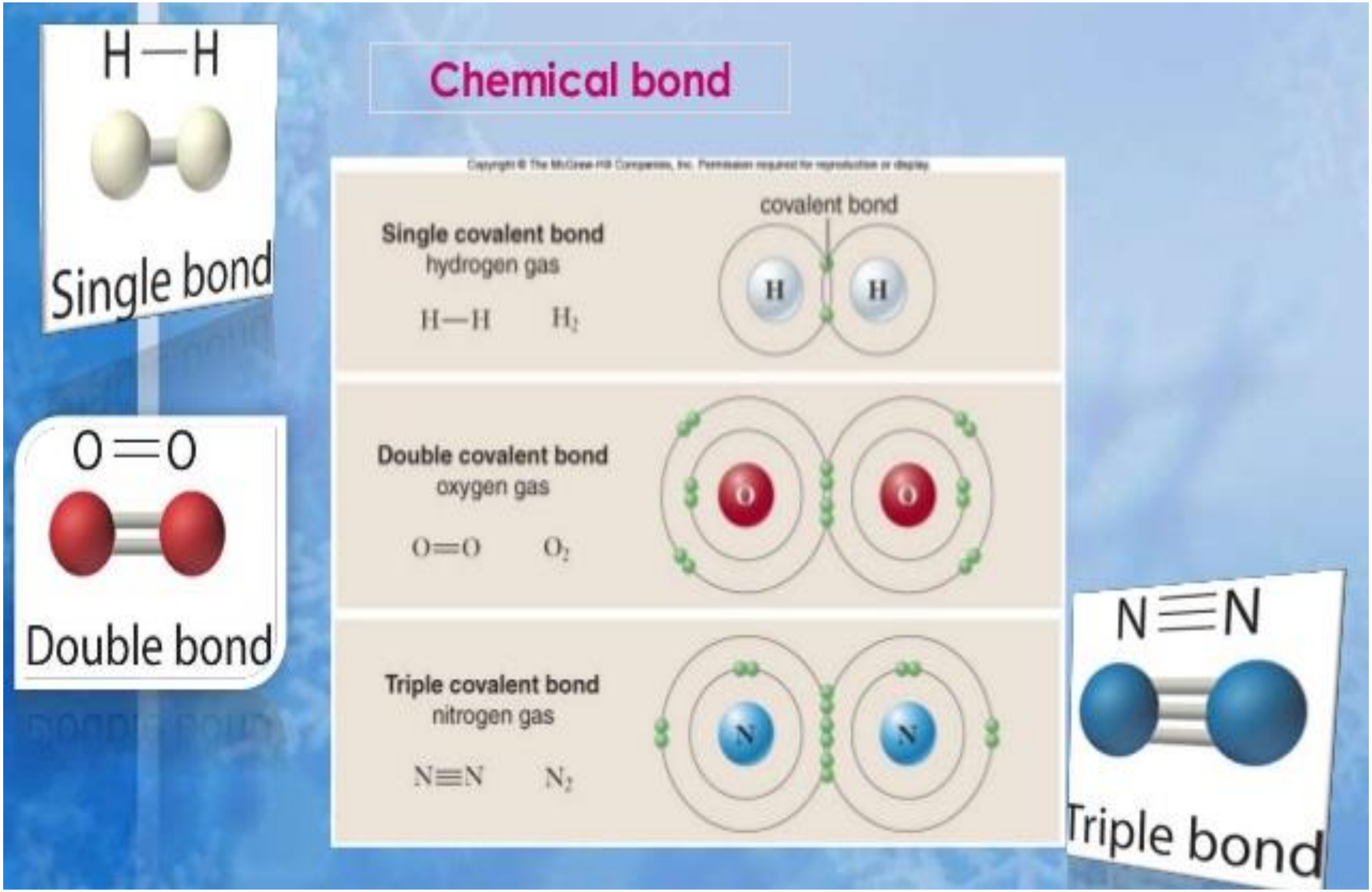
- Crystalline solids
- High melting and boiling points
- Conduct electricity when melted
- Many soluble in water but not in nonpolar liquid

Properties of Covalent Compounds

- Gases, liquids, or solids
- Low melting and boiling points
- Poor electrical conductors in all phases
- Many soluble in non polar liquids but not in water
- **Are brittle**
- **When 2 atoms bond covalently the resulting particle is a molecule**

What Have You Learned?

- What are the two extreme cases of bonds?
- Do covalent bonds lose or gain electrons? (Yes or No).
- Why do atoms bond?


Chemical bond
H—H
Single bond
O=O
Double bond
N≡N
Triple bond
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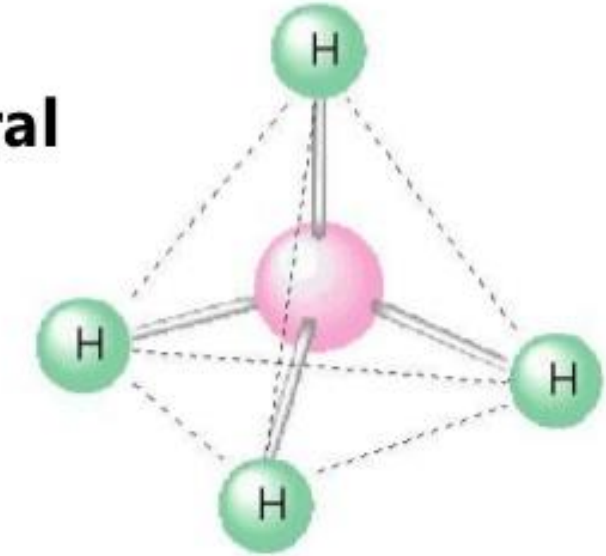
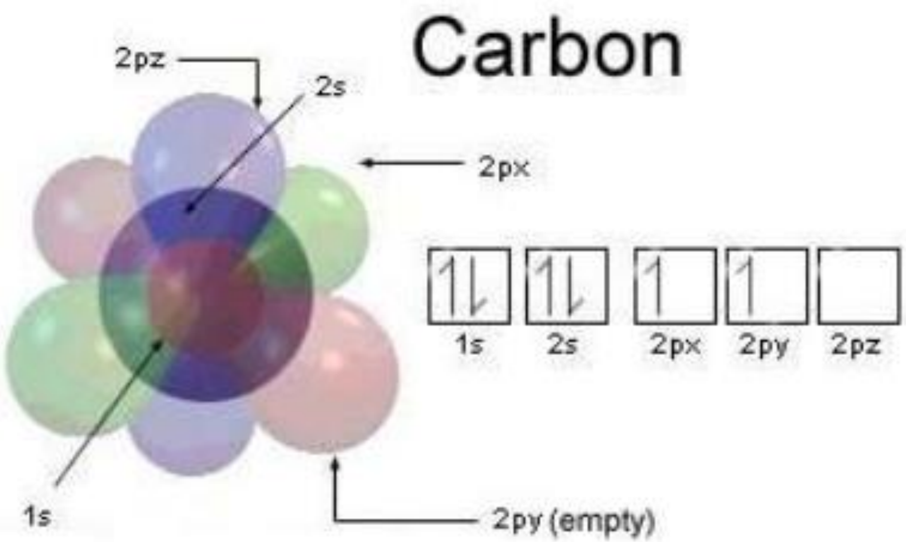
Single covalent bond
hydrogen gas
H—H
H2
covalent bond
H
H
Double covalent bond
oxygen gas
O=O
O2
O
O
Triple covalent bond
nitrogen gas
N≡N
N2
N
N

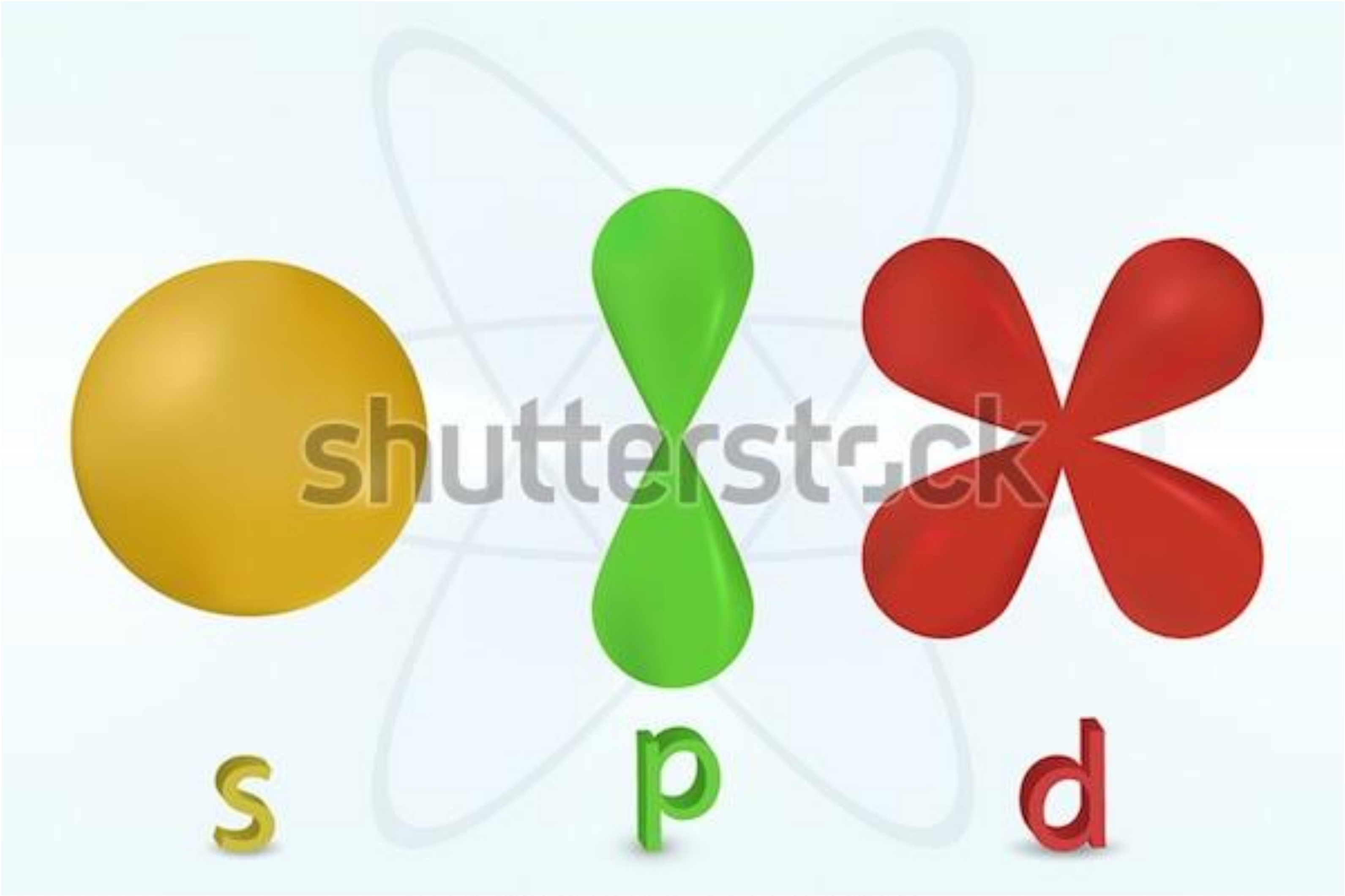
Recall : methane, CH_4 – tetrahedral shape (4 equal covalent bonds)



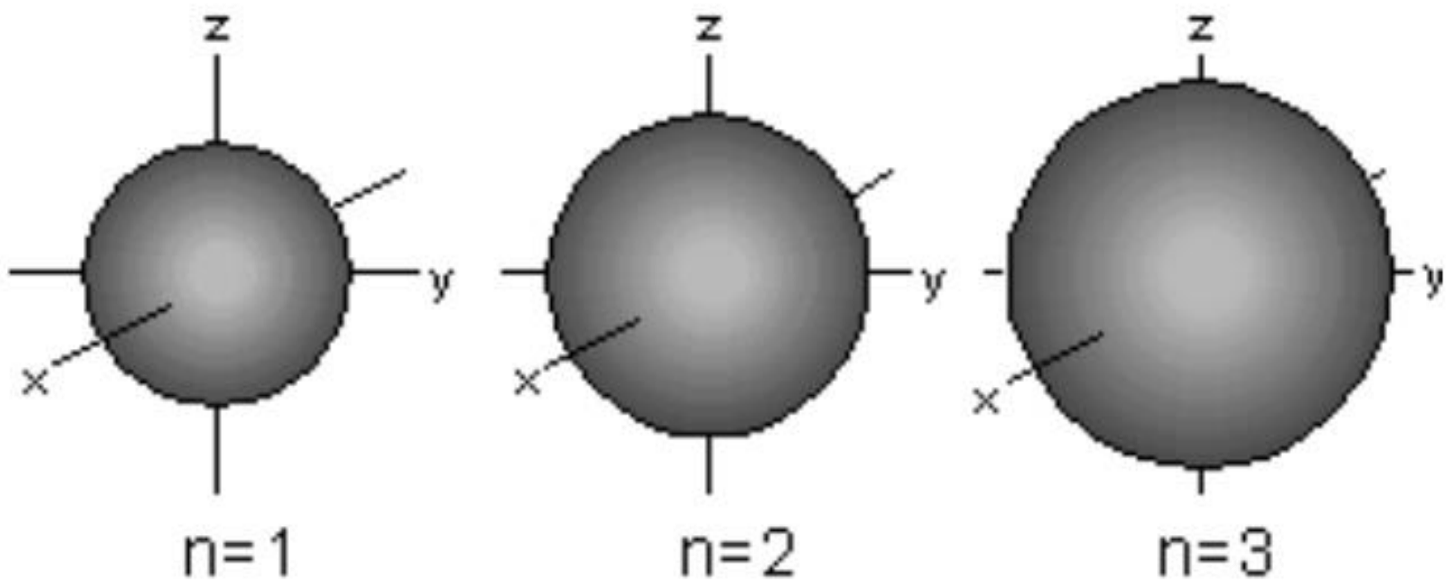


Carbon has an electron arrangement $1s^2 2s^2 2p^2$

There is a contradiction here . . . . . .

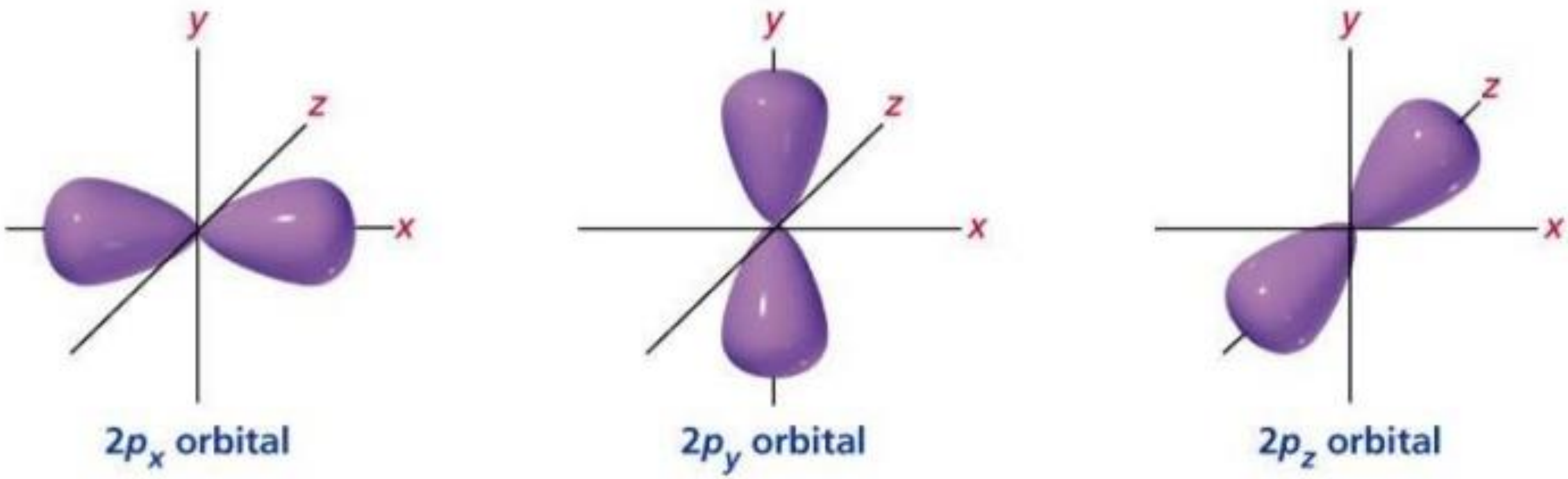

s
p
d

Recall: the s orbital shape

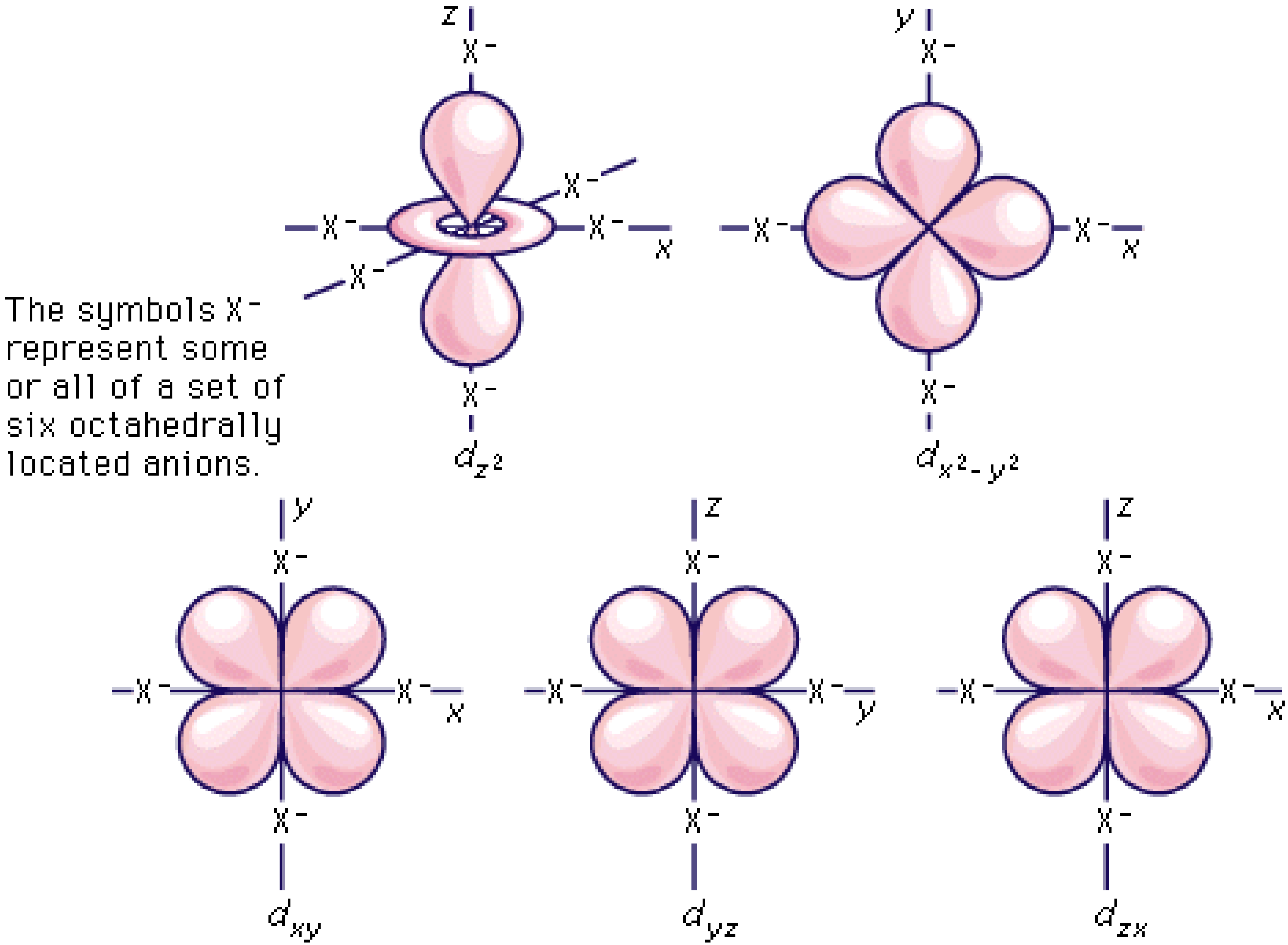


ONE s orbital in each energy level

Recall: the p orbital shape

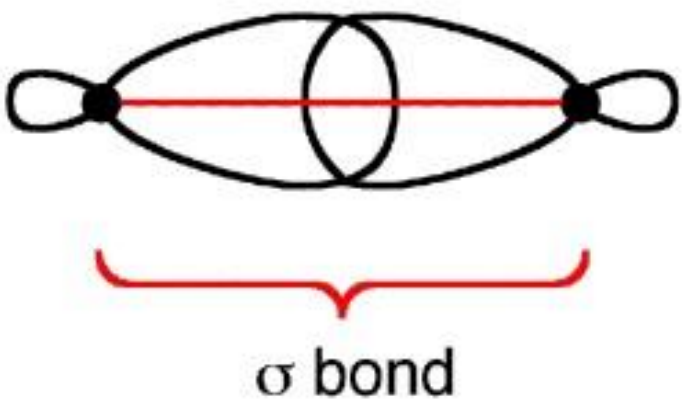


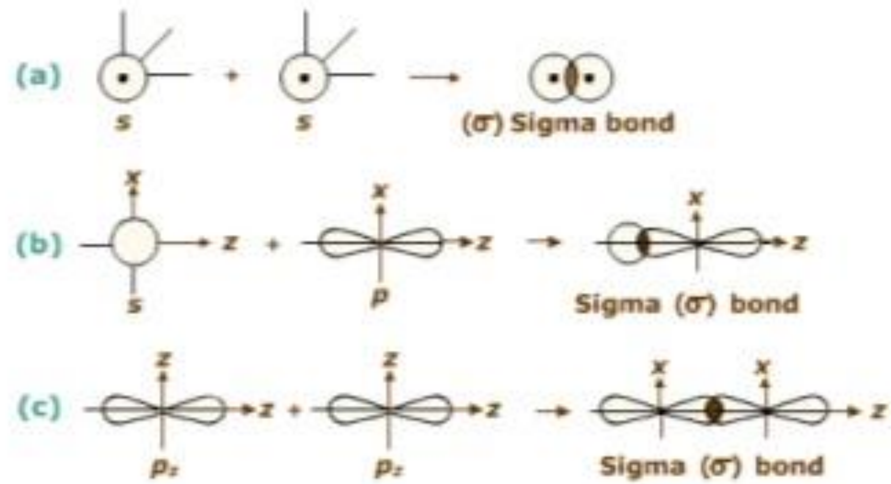
THREE p orbitals in each energy level (beginning n = 2)


The symbols X⁻ represent some or all of a set of six octahedrally located anions.
z
X⁻
x
y
d z²
d x²-y²
d xy
d yz
d zx

Sigma bond

In chemistry, **sigma bonds** (**σ bonds**) are the strongest type of covalent chemical bond. They are formed by head-on overlapping between atomic orbitals. Sigma bonding is most clearly defined for diatomic molecules.

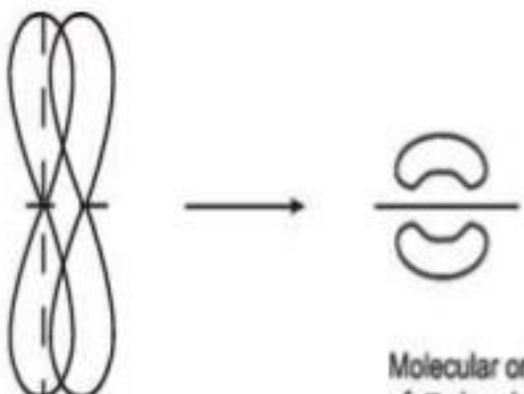




Formation of a sigma bond due to (a) The s - s overlap (b) The s - p overlap (c) The p_z - p_z overlap

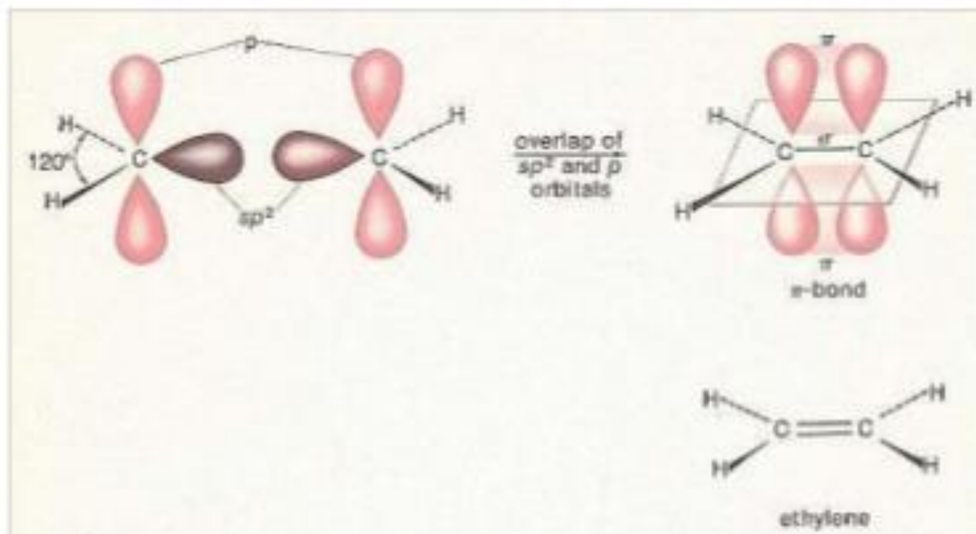
π bond

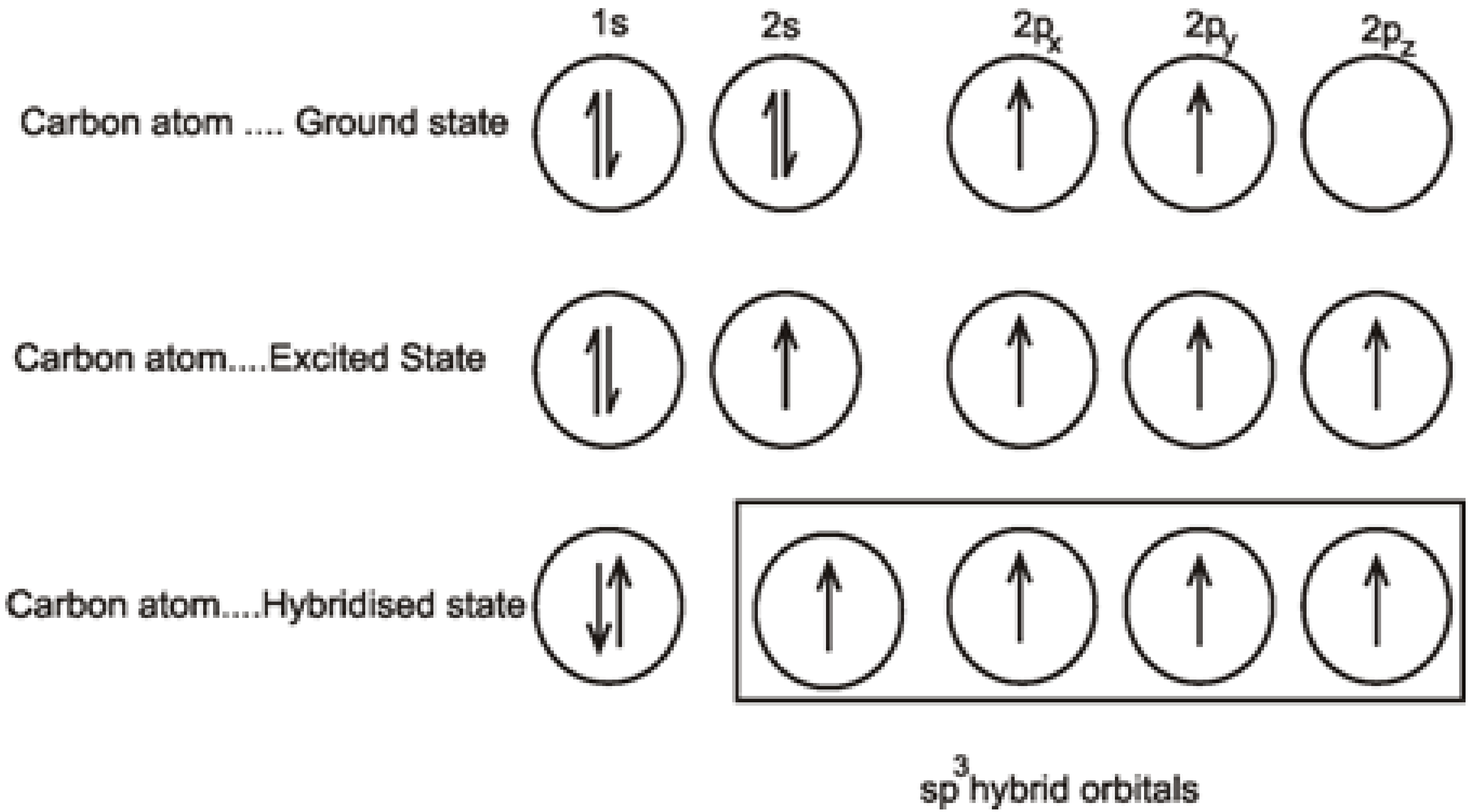
Pi bonds **(π bonds)** are covalent chemical bonds where two lobes of one involved atomic orbital overlap two lobes of the other involved atomic orbital. Each of these atomic orbitals is zero at a shared nodal plane, passing through the two bonded nuclei.



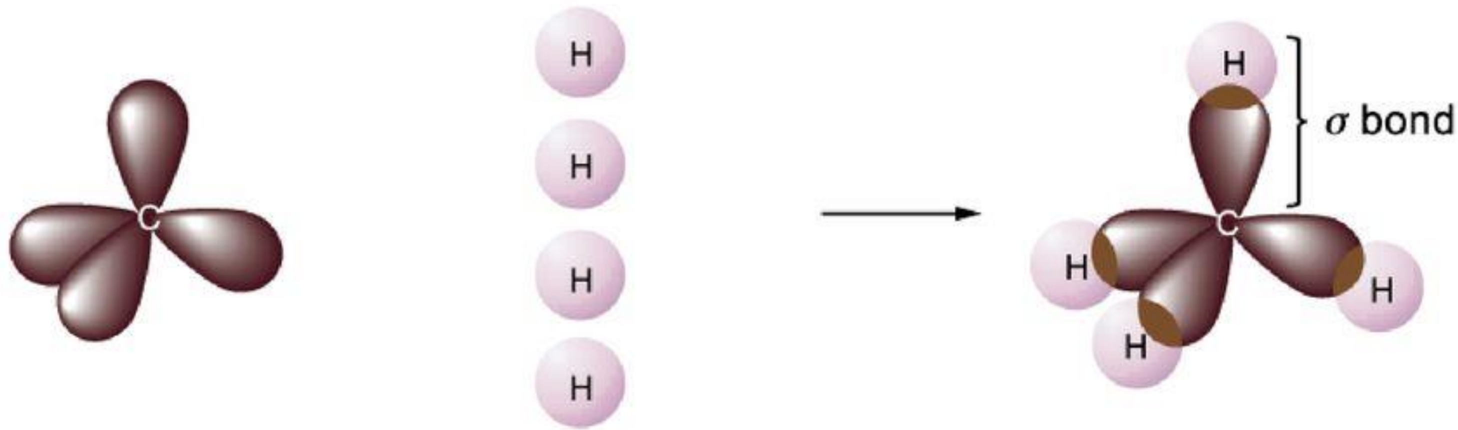
Sidewise overlap of p - orbital

Molecular orbital of π - bond




1s
2s
2px
2py
2pz
Carbon atom .... Ground state
Carbon atom....Excited State
Carbon atom....Hybridised state
sp3 hybrid orbitals

Methane (CH_4)

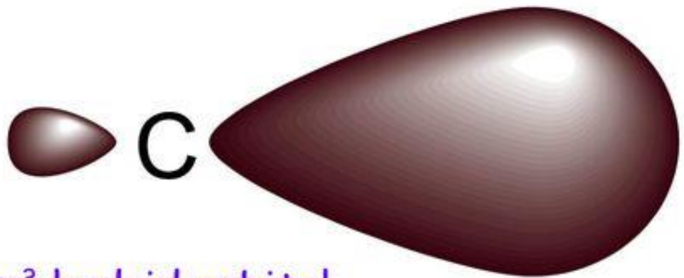


Four sp^3 hybrid orbitals of a carbon atom

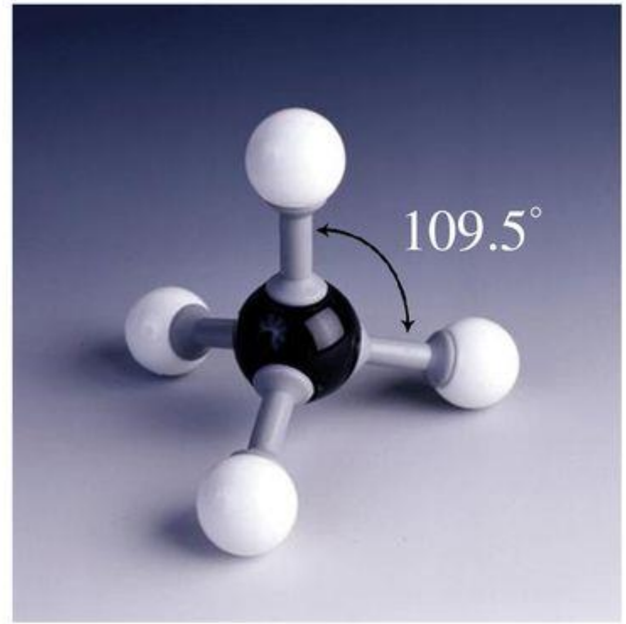
Four 1*s* orbitals of four hydrogen atoms

Overlapping of orbitals in a methane molecule

New Shape for New Orbitals



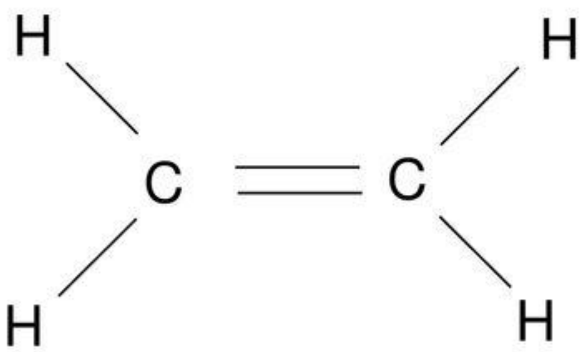
sp^3 hydrid orbital

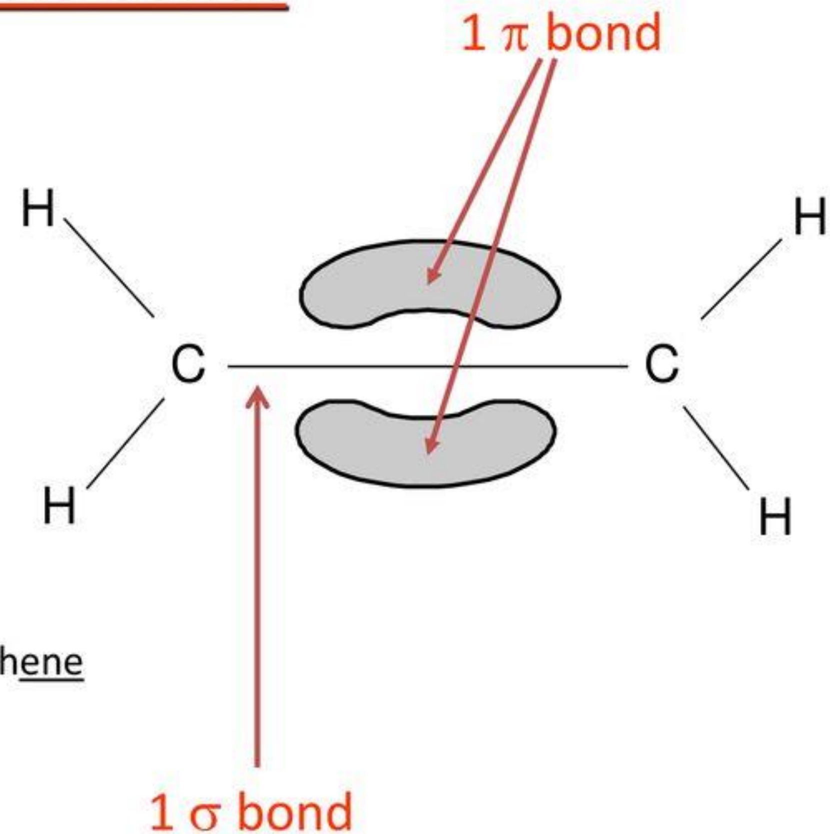


Head-on **overlapping** of **orbitals**

Alkenes Contain Carbon-Carbon

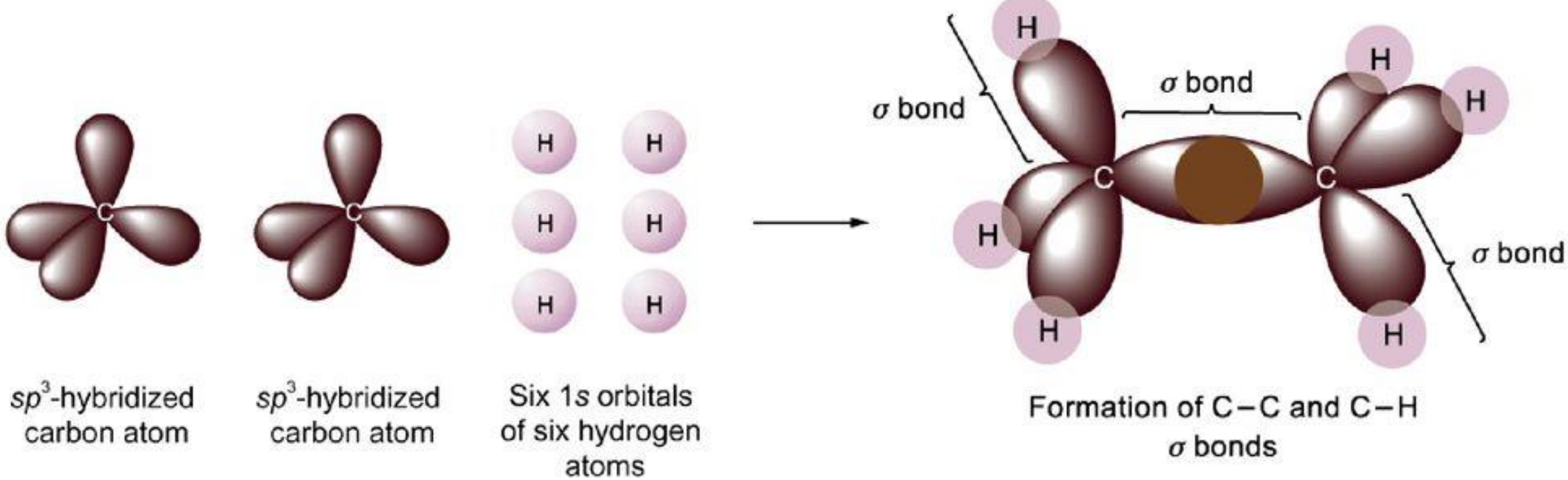
Double bonds

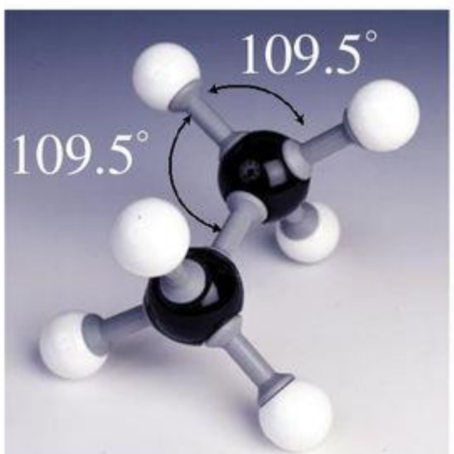




Ethene

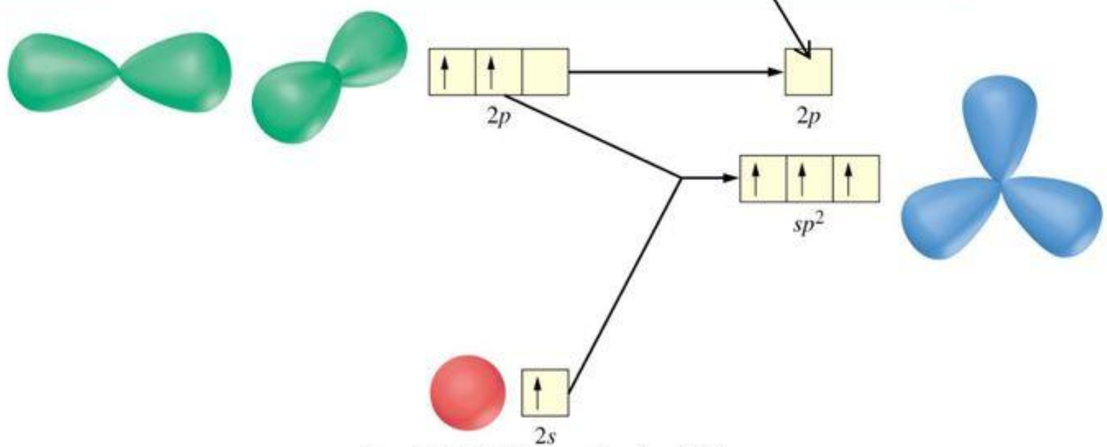
Ethane (C_2H_6)

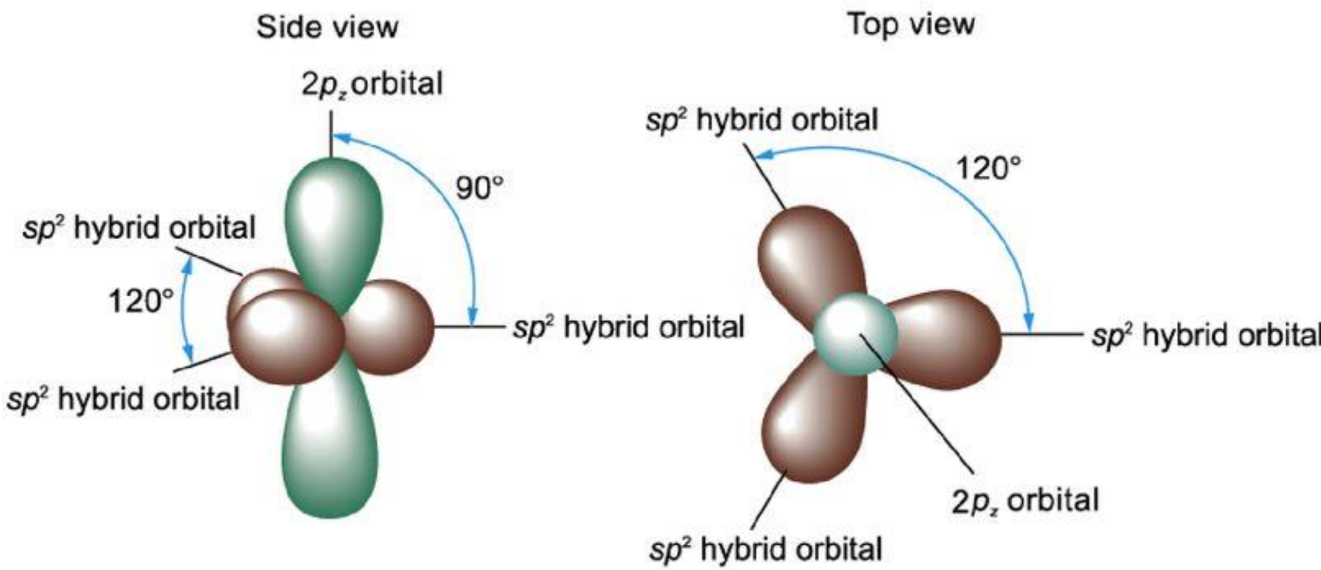

C
C
H
sp3-hybridized carbon atom
sp3-hybridized carbon atom
Six 1s orbitals of six hydrogen atoms
σ bond
σ bond
σ bond
Formation of C–C and C–H σ bonds

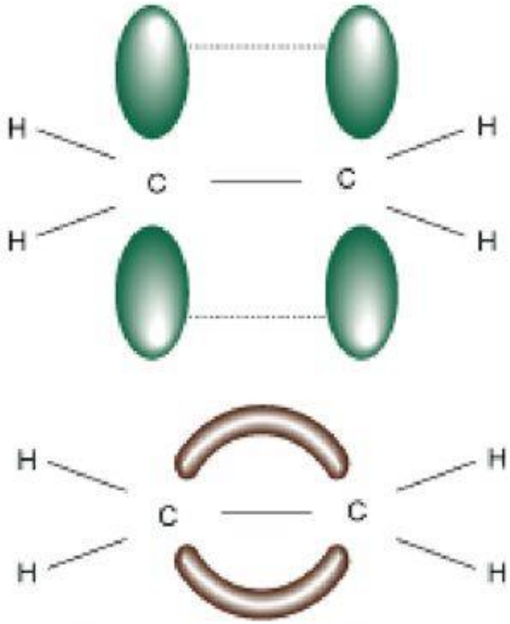

109.5°
109.5°

sp^2 *Hybridization*

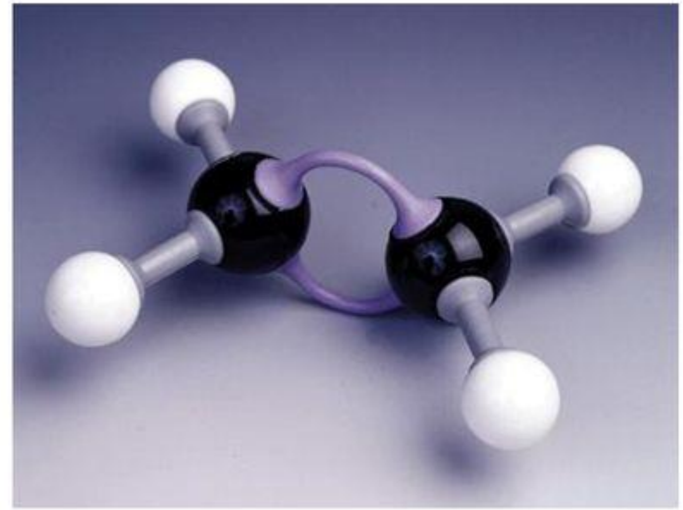
Unhybridized ***p***

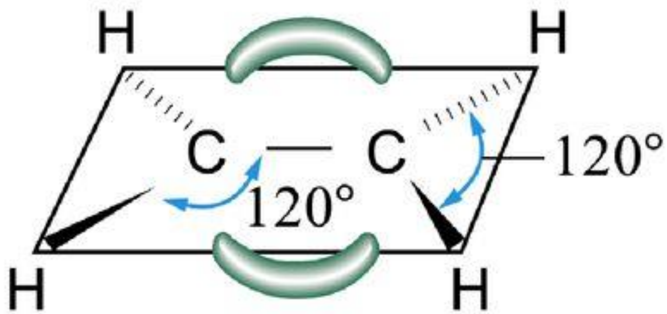




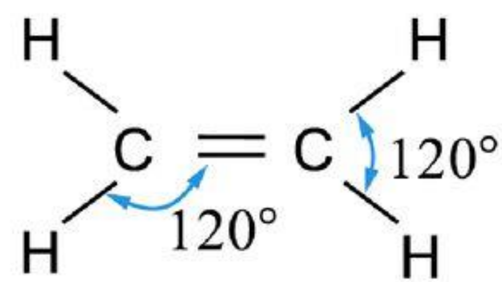


Formation of π bond by the side-way overlap of $2p_z$ orbitals





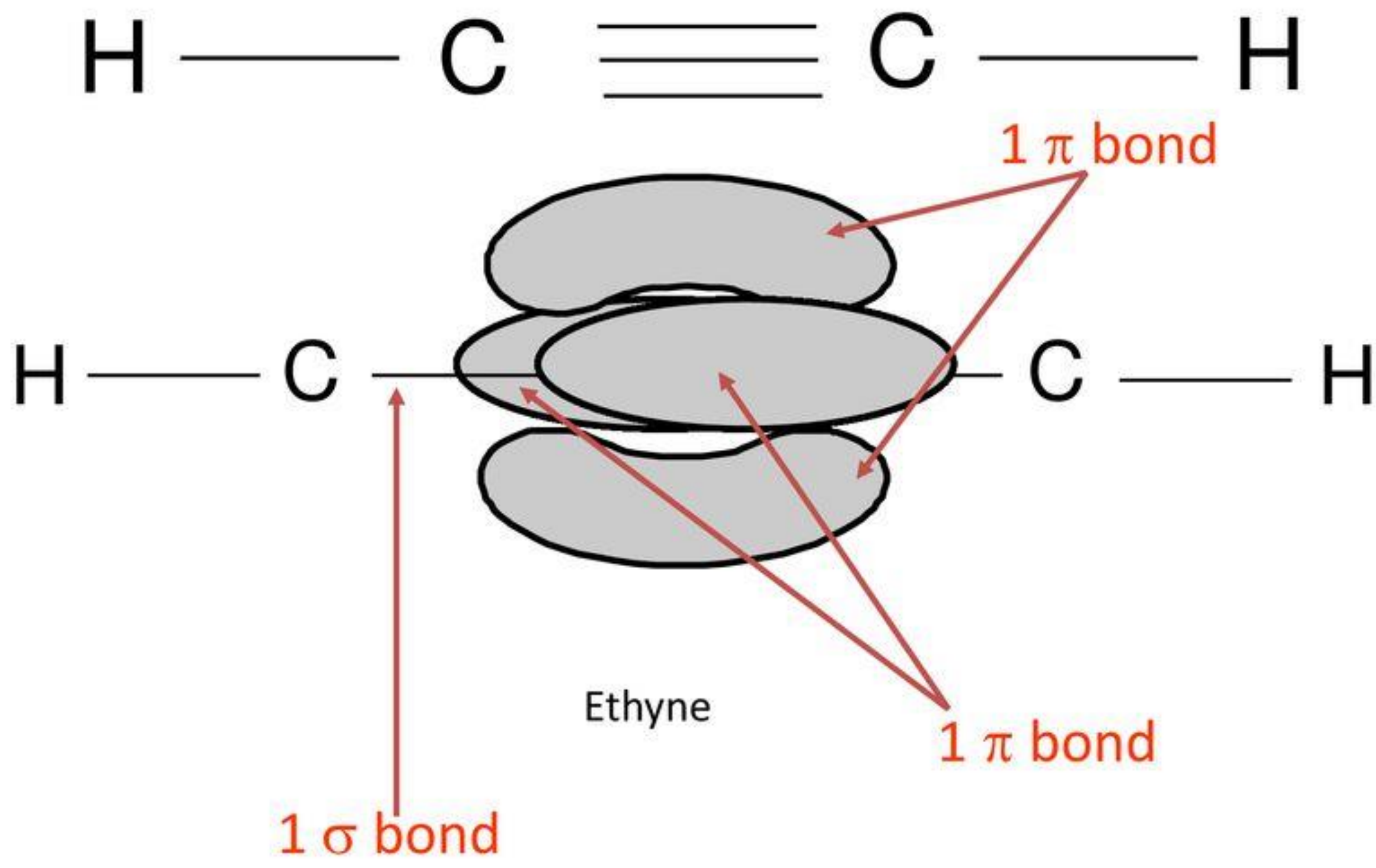
or



(a) Differences between the σ bond and the π bond in the carbon-carbon double bond:

σ bond	**π bond**
Head-on overlap of the *sp*^{2} hydridized orbitals of two carbon atoms	Side-way overlap of the vacant *p* orbitals of two carbon atoms
The bonding electrons in are localized symmetrically along the internuclear axis	The electrons in the π bond appear as two lobes, one above and one below the internuclear axis
Stronger	Weaker
Free to rotate	Restricted to rotation

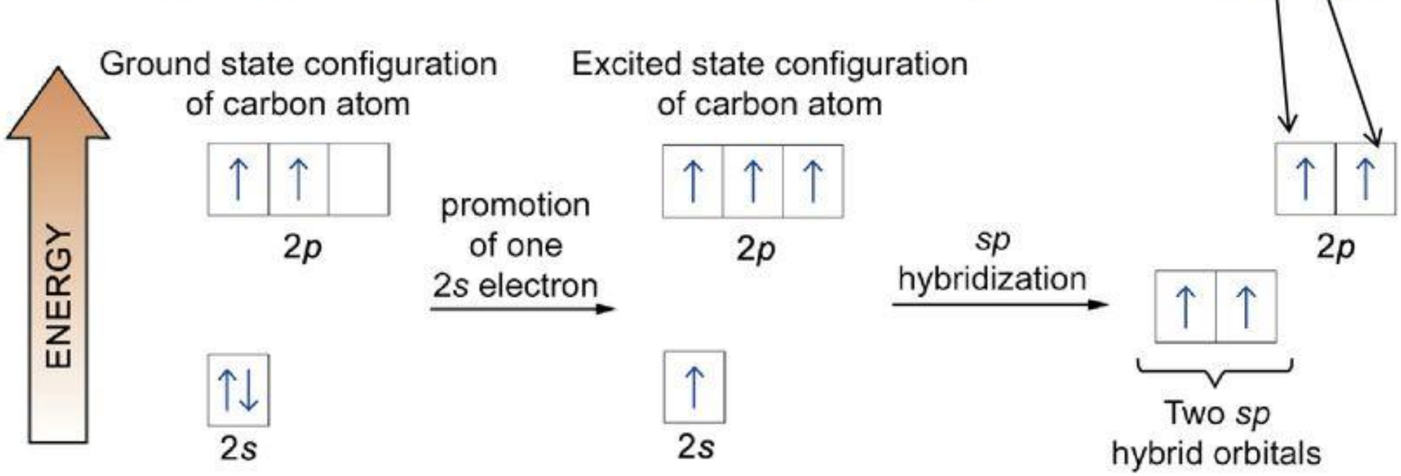
Alkynes Contain Carbon-Carbon Triple Bonds

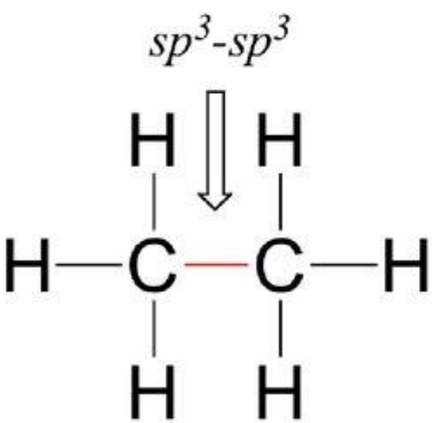


In

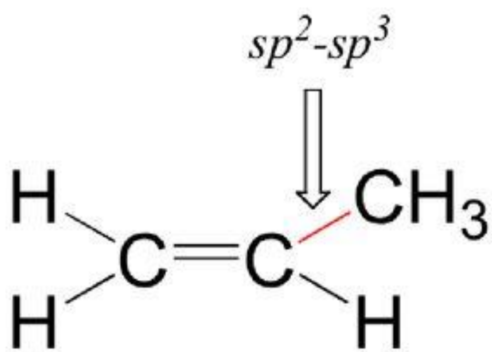
sp Hybridization

2 Unhybridized ***p for pi bonds***

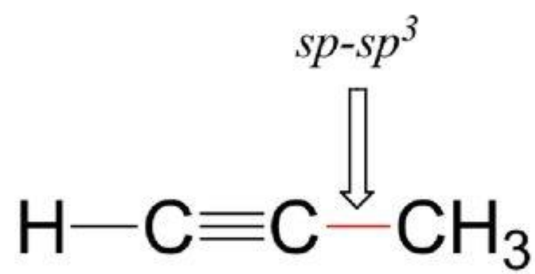




A



B



C

PRINCIPLES OF MOT

- In molecules, atomic orbitals combine to form molecular orbitals which surround the molecule.

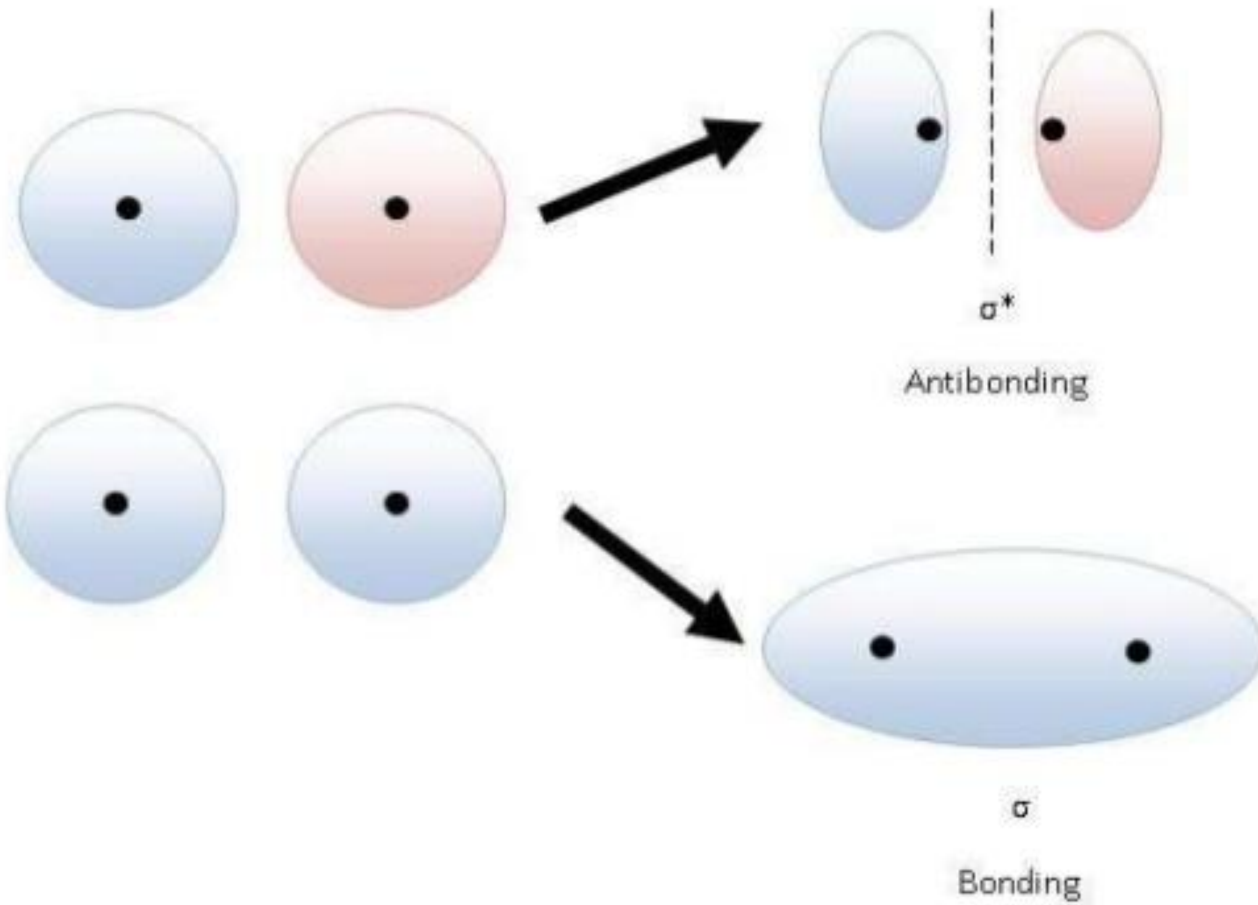
Molecular Orbitals

Bonding　　　Antibonding

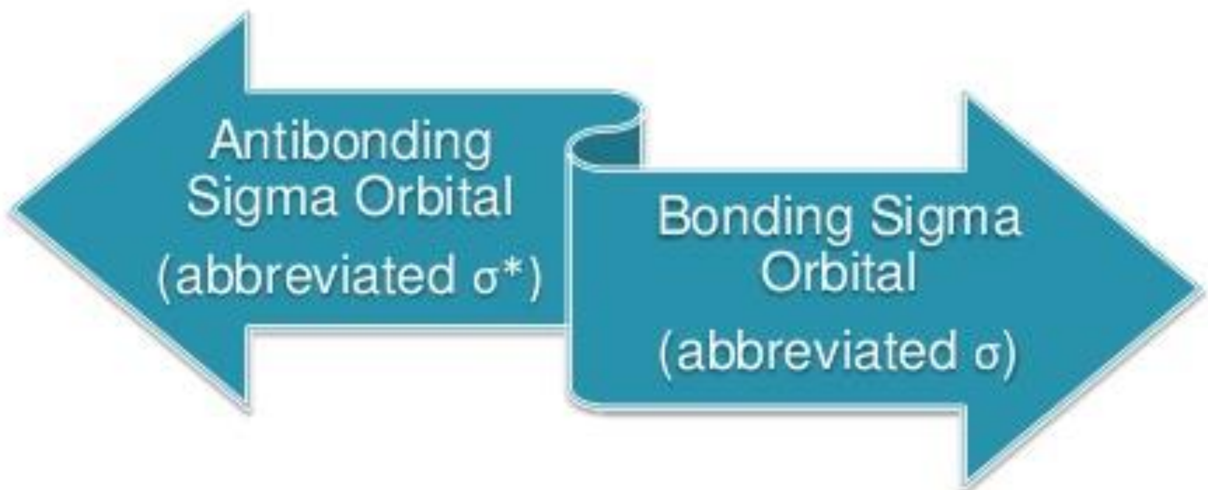
- Molecular bonds have lower potential energy than in separate atomic orbitals.
- Thus, electrons prefer to stay in a molecular bond.

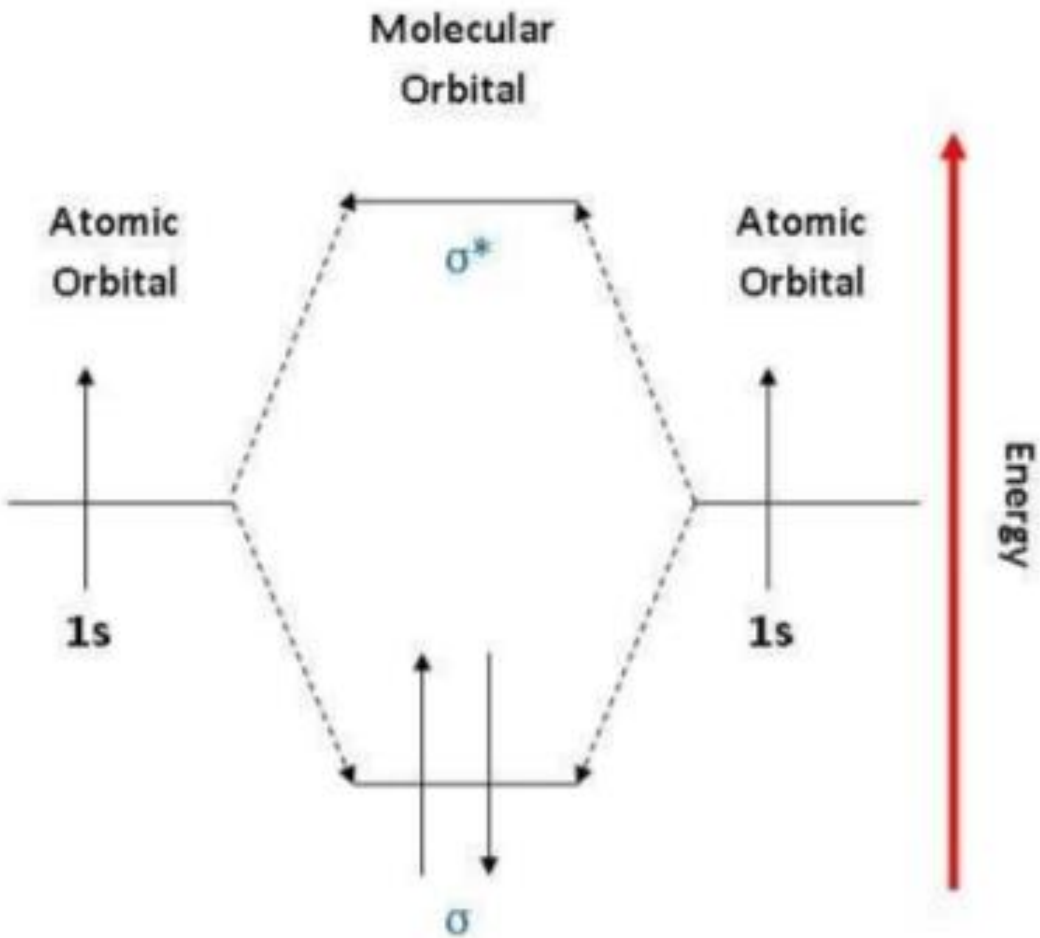
Antibonding
Molecular
Orbital

Bonding
Molecular
Orbital


σ*
Antibonding
σ
Bonding

SIGMA BONDS (σ)

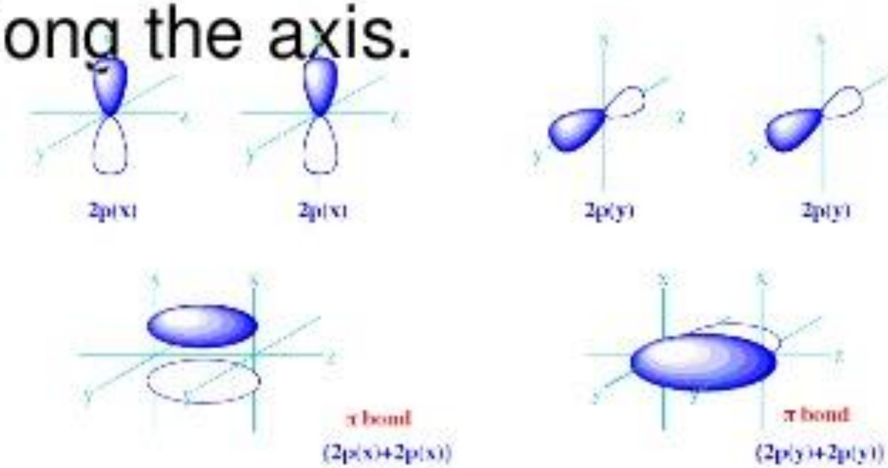
- They are symmetrical about the axis
- 



Molecular
Orbital
σ*
Atomic
Orbital
Atomic
Orbital
1s
1s
σ
Energy

PI BONDS (π)

- The pi bonding bonds as a side to side overlap, which then causes there to be no electron density along the axis, but there is density above and belong the axis.



DETERMINING BOND ORDER

- Bond Order indicates the strength of the bond. The higher the Bond Order, the stronger the bond.

$$\text{Bond Order} = 1/2(a-b)$$

where, a = number of e^- in bonding Molecular Orbitals

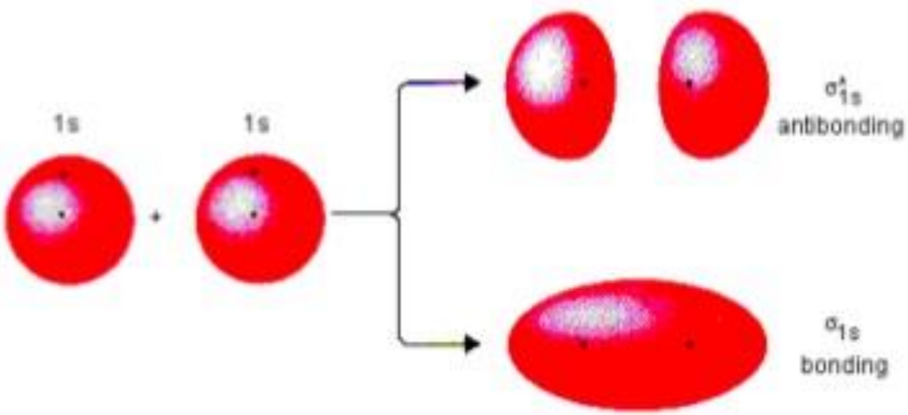
b = number of e^- in antibondng Molecular Orbitals.

STABILITY OF THE MOLECULE

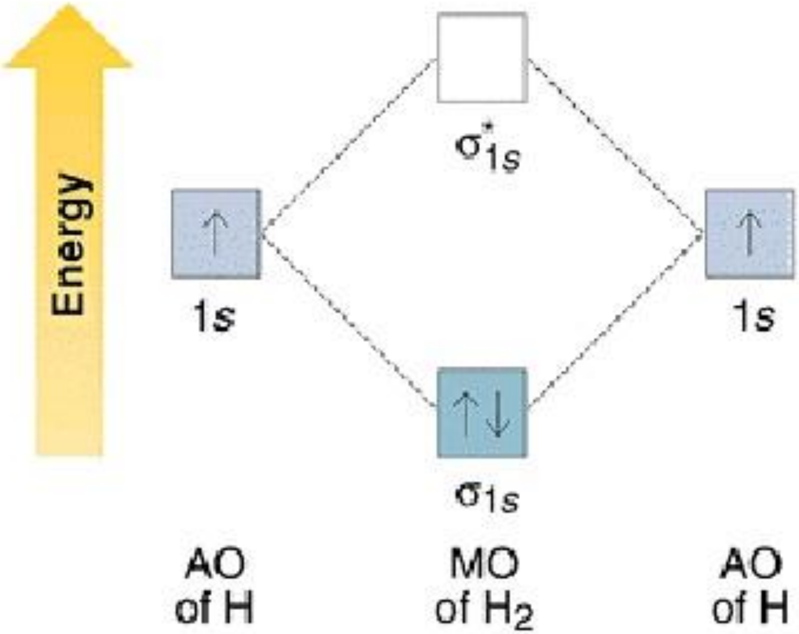
- If the Bond Order is Zero, then no bonds are produced and the molecule is not stable (for example He_2).
- If the Bond Order is 1, then it is a single covalent bond.
- The higher the Bond Order, the more stable the molecule is.

Molecular Orbital Theory

- Molecular orbitals result from the combination of atomic orbitals. Since orbitals are wave functions, they can combine either constructively (forming a bonding molecular orbital), or destructively (forming an antibonding molecular orbital).
- Consider the H_2 molecule, for example. One of the molecular orbitals in this molecule is constructed by adding the mathematical functions for the two 1*s* atomic orbitals that come together to form this molecule. Another orbital is formed by subtracting one of these functions from the other



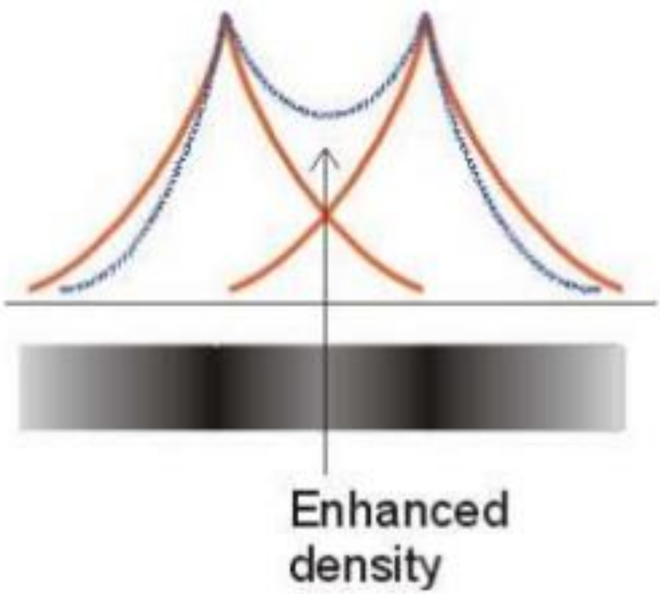
Electronic configuration, σ^2, for H_2 molecules



H_2 bond order $= \frac{1}{2}(2 - 0) = 1$

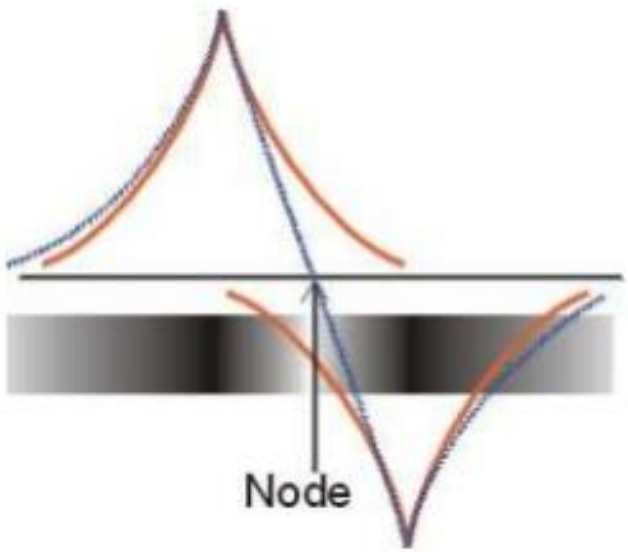
Generalize the technique of LCAO

Bonding Molecular Orbital Theory



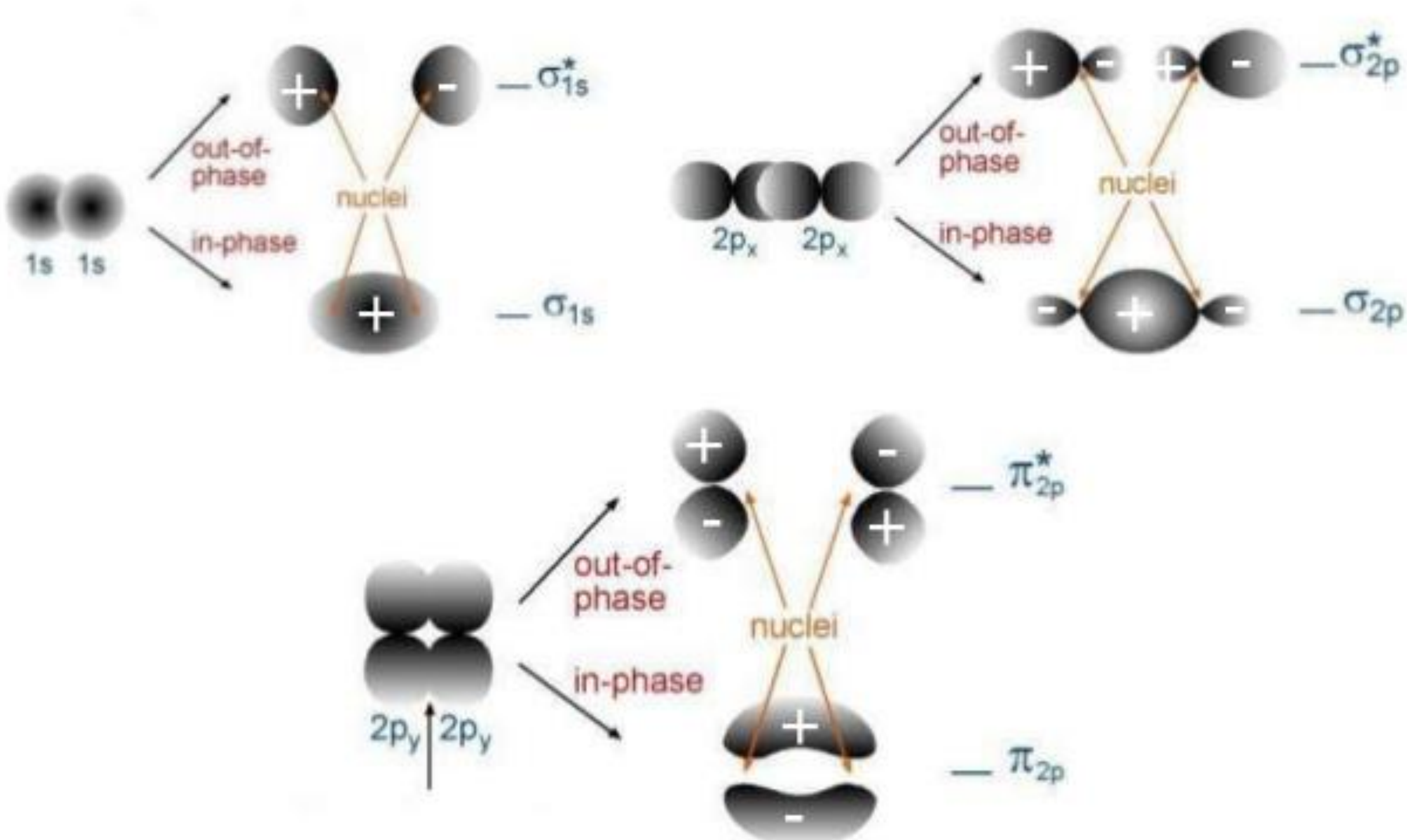
The bonding orbital results in increased electron density between the two nuclei, and is of lower energy than the two separate atomic orbitals.

Antibonding Molecular Orbital Theory



The antibonding orbital results in a node between the two nuclei, and is of greater energy than the two separate atomic orbitals.

Overlap of s & p Orbitals



Sigma bonding orbitals

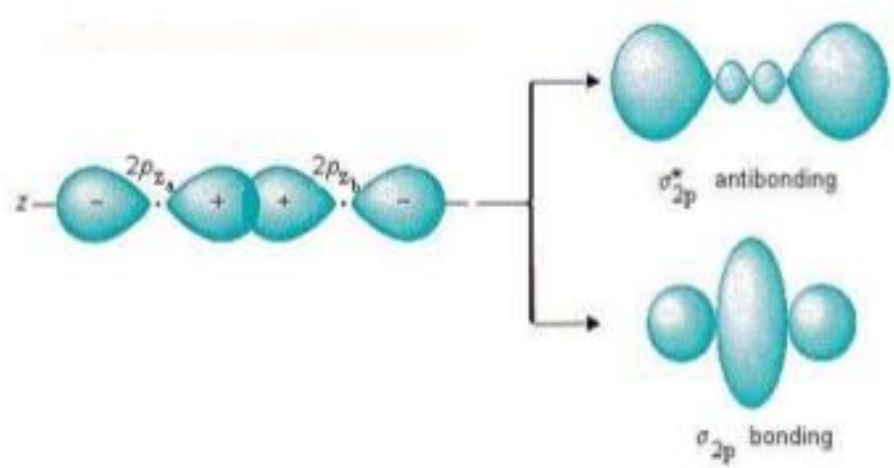
- From s orbitals on separate atoms

s orbital s orbital

Sigma bonding
molecular orbital

Molecular Orbitals of the Second Energy Level

If we arbitrarily define the *Z* axis of the coordinate system for the O_2 molecule as the axis along which the bond forms, the $2p_z$ orbitals on the adjacent atoms will meet head-on to form a $_{2p}$ bonding and a $_{2p}$* antibonding molecular orbital

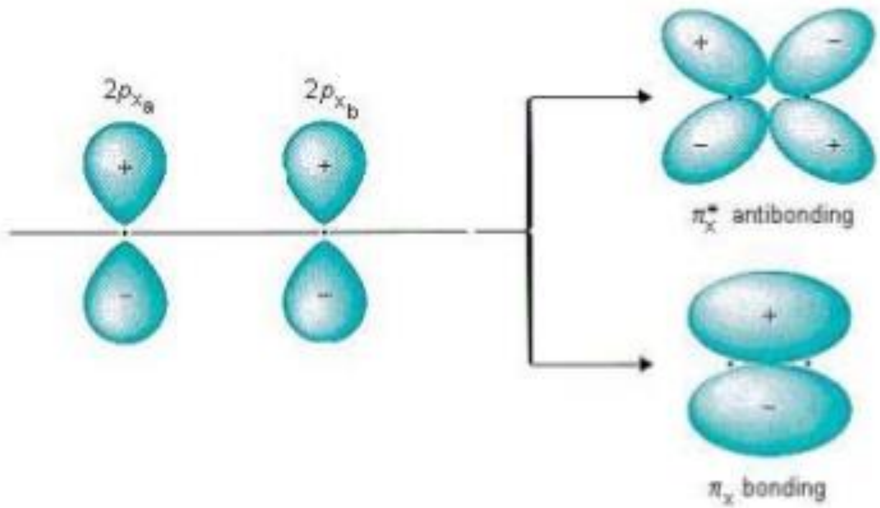


Sigma bonding orbitals

- From p orbitals on separate atoms

p orbital p orbital

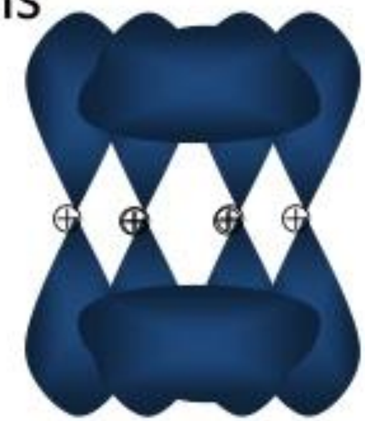
Sigma bonding
molecular orbital



The $2p_x$ orbitals on one atom interact with the $2p_x$ orbitals on the other to form molecular orbitals that have a different shape. These molecular orbitals are called *pi* (π) orbitals because they look like *p* orbitals when viewed along the bond.

Pi bonding orbitals

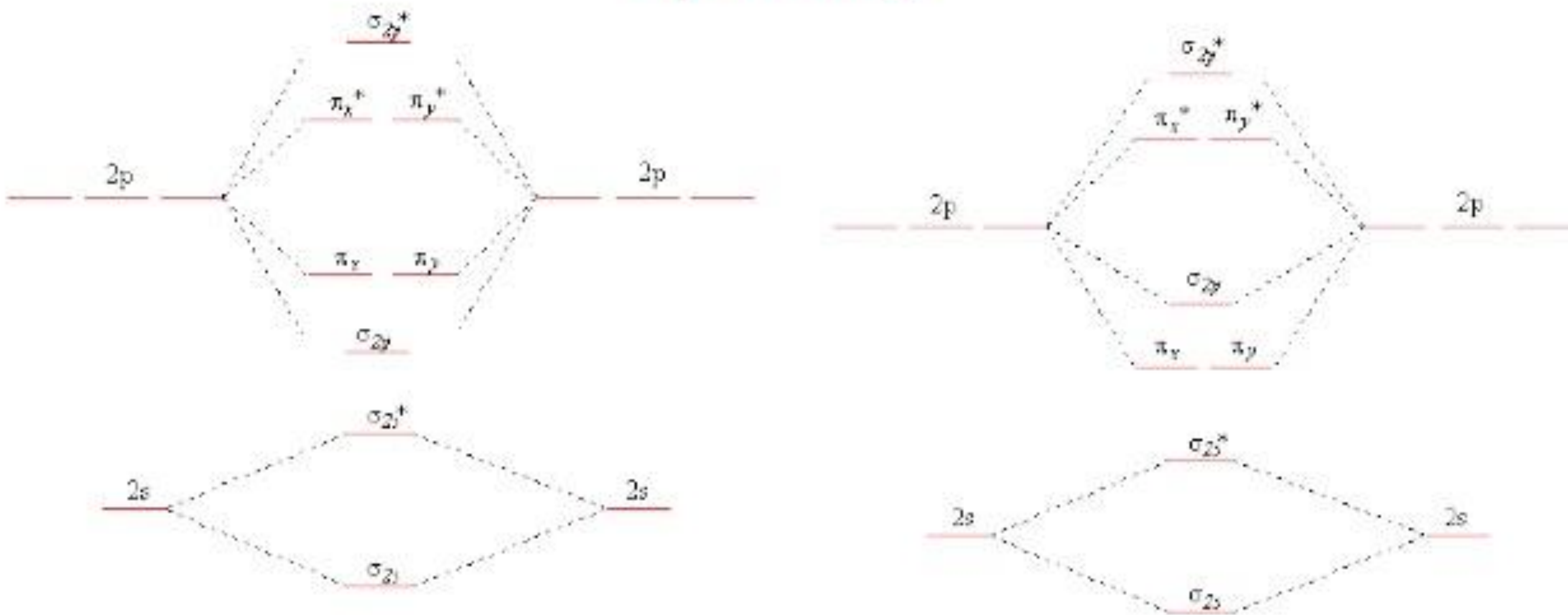
- P orbitals on separate atoms



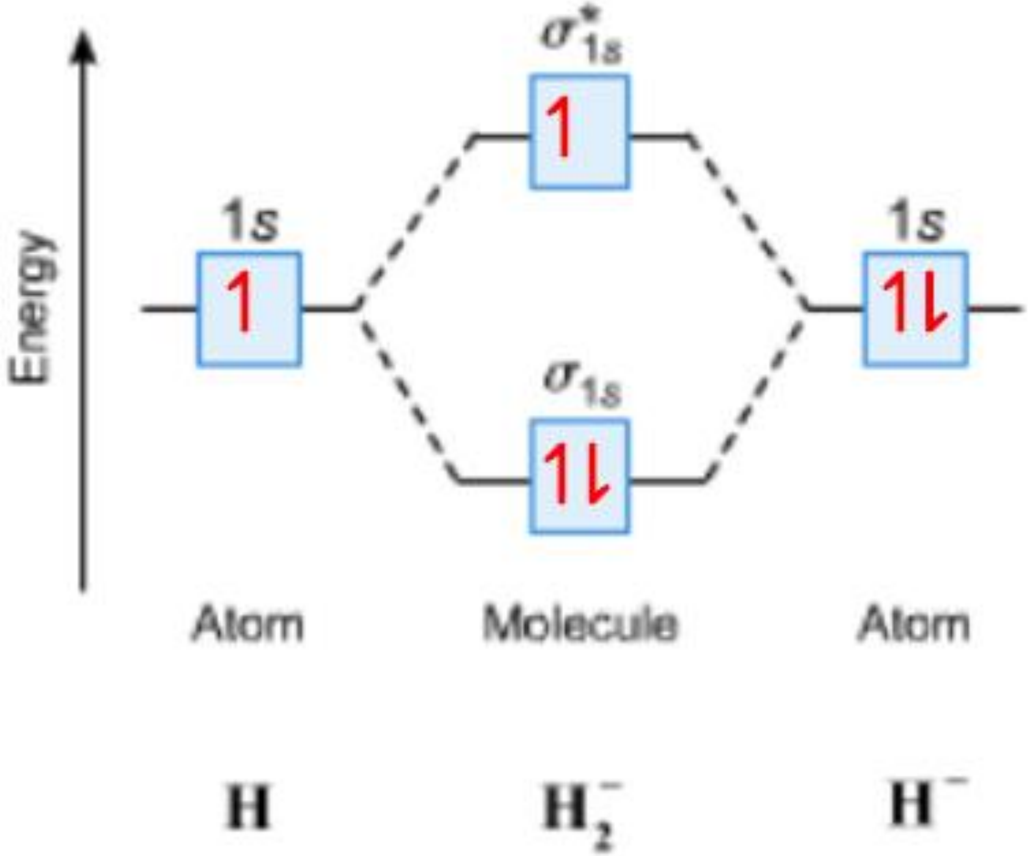
Pi bonding
molecular orbital

Molecular Orbital Diagram

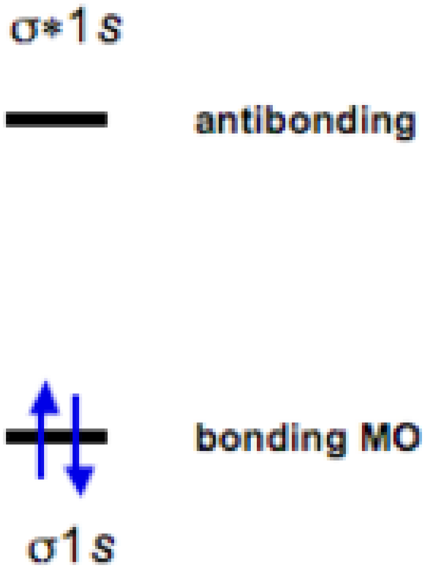
s-p mixing



MOLECULAR ORBITAL OF H2 MOLECULE



H_2



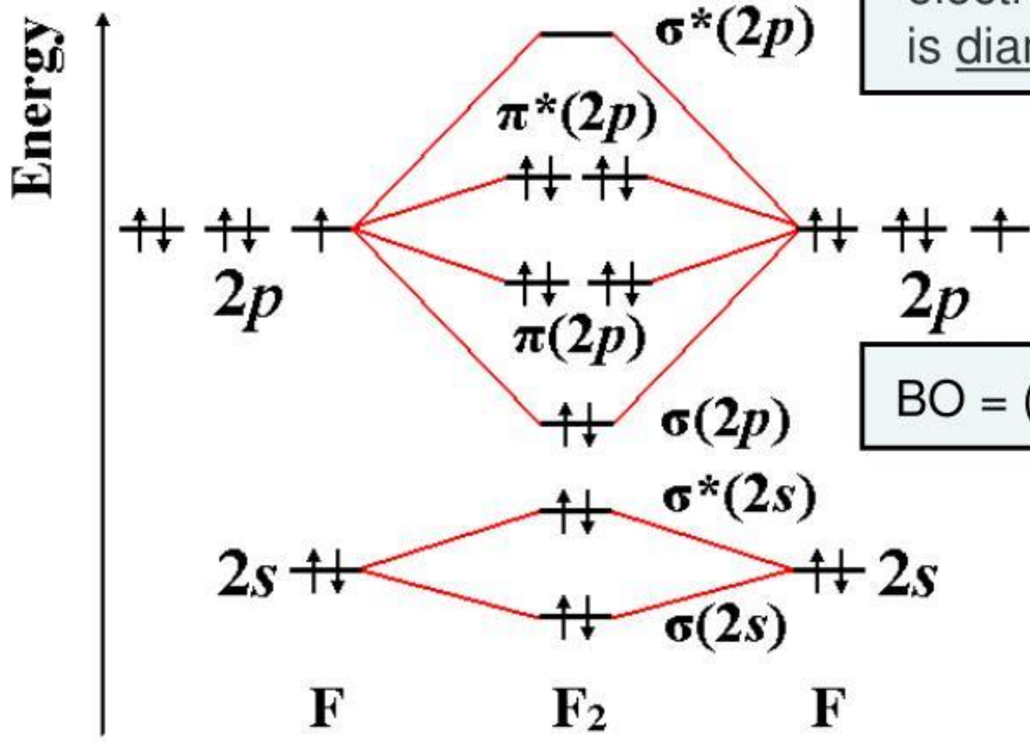
$$\text{bond order} = \frac{\begin{pmatrix}\text{number of electrons}\\ \text{in bonding MO}\end{pmatrix} - \begin{pmatrix}\text{number of electrons}\\ \text{in anti-bonding MO}\end{pmatrix}}{2}$$

$$\frac{2 - 0}{2} = 1$$

bond order = 1

F_2 molecule, bond order = 1

F_2 has no unpaired electrons, and so is diamagnetic



BO = (6-4)/2 = 1

MO diagram for difluorine (F_2) molecule.

Thank you

PARTIAL IONIC CHARACTER

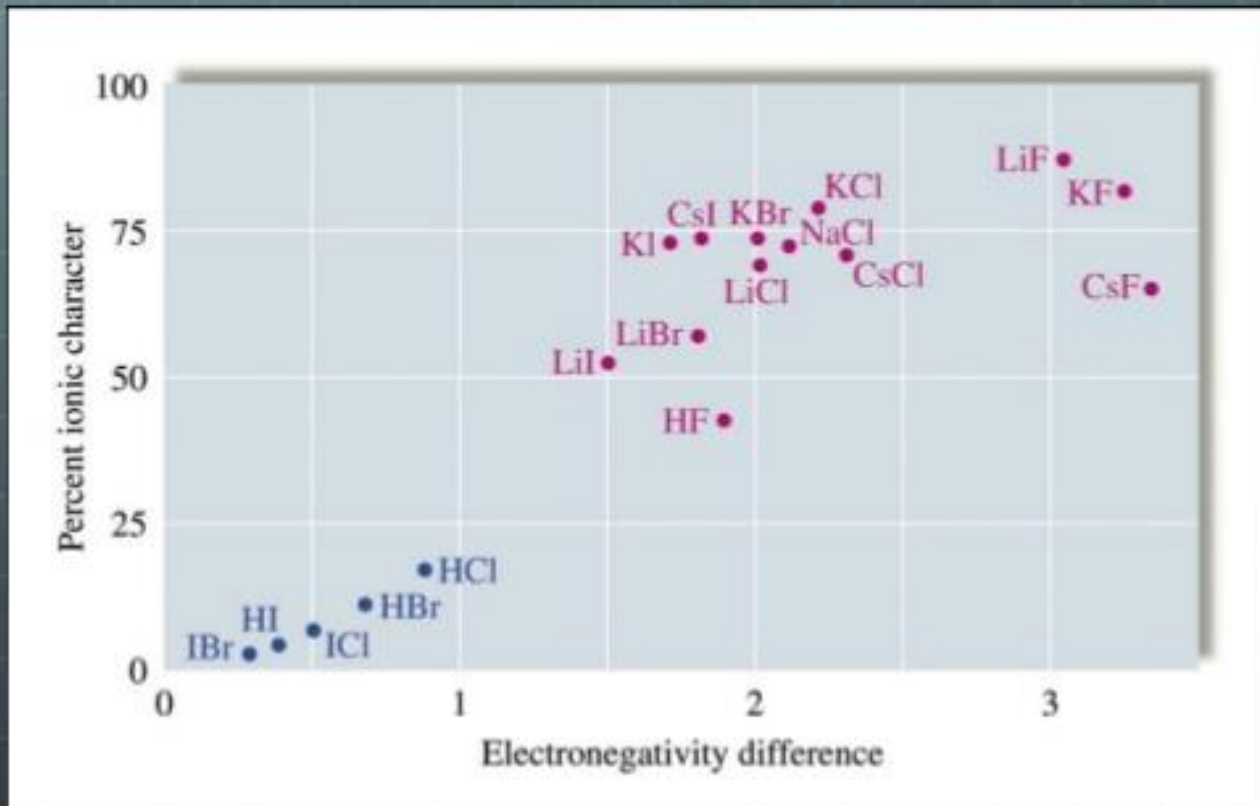
Dipoles and Dipole Moments

- Electrostatic potential diagram shows variation in charge. Red is the most electron rich region and blue is the most electron poor region.

Dipoles and Dipole Moments

- Dipole moments are when opposing bond polarities don't cancel out.

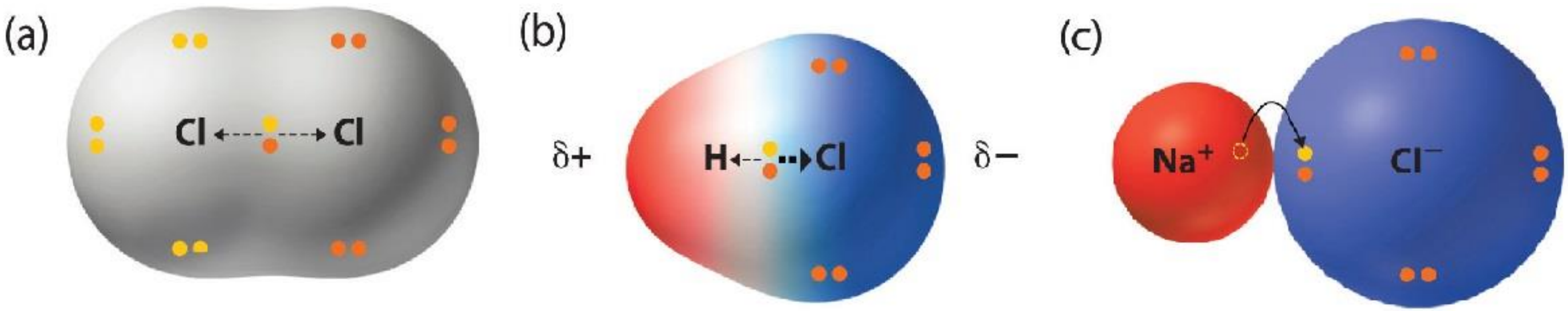
Bond Character



Bond Energy

TABLE 8.5 Bond Lengths for Selected Bonds

Bond	Bond Type	Bond Length (pm)	Bond Energy (kJ/mol)
C—C	Single	154	347
C═C	Double	134	614
C≡C	Triple	120	839
C—O	Single	143	358
C═O	Double	123	745
C—N	Single	143	305
C═N	Double	138	615
C≡N	Triple	116	891



Nonpolar covalent bond

Bonding electrons shared equally between two atoms. No charges on atoms.

Polar covalent bond

Bonding electrons shared unequally between two atoms. Partial charges on atoms.

Ionic bond

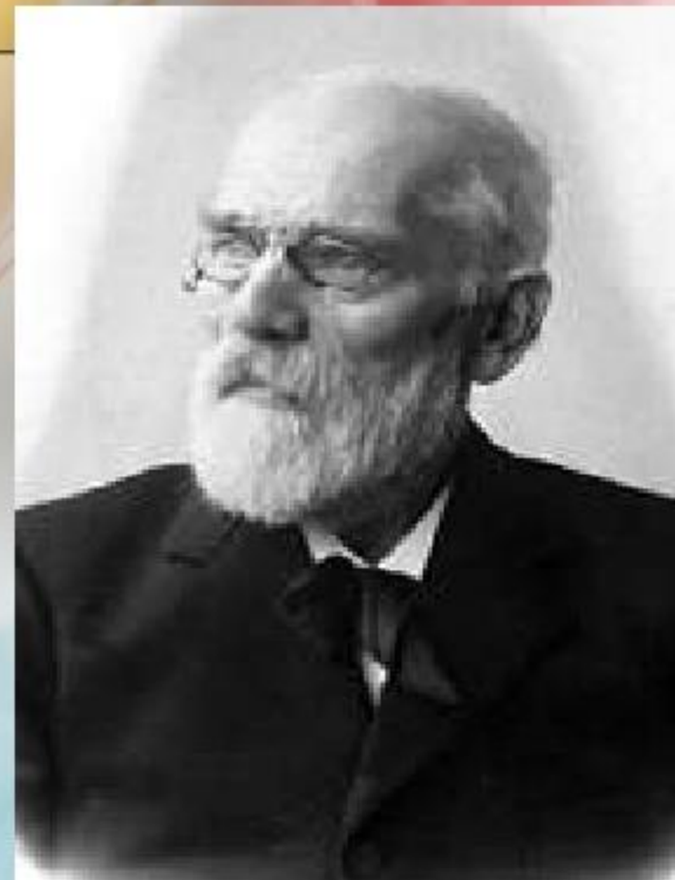
Complete transfer of one or more valence electrons. Full charges on resulting ions.

- Why do some solids dissolve in water but others do not?
- What happens when noble gases liquefy?
- Why are some substances gases at room temperature, but others are liquid or solid?

The answers lie in………

Intermolecular Forces

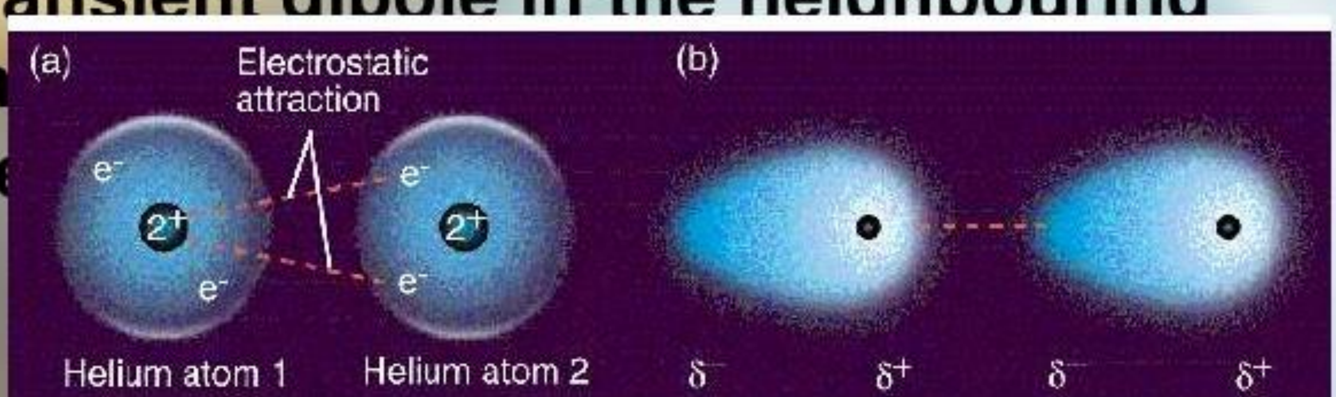
- **Intermolecular forces** may be attractive or repulsive.
- **Johannes D van der Waals**, Dutch, was the first to postulate intermolecular forces in developing a theory to account for properties of real gases.



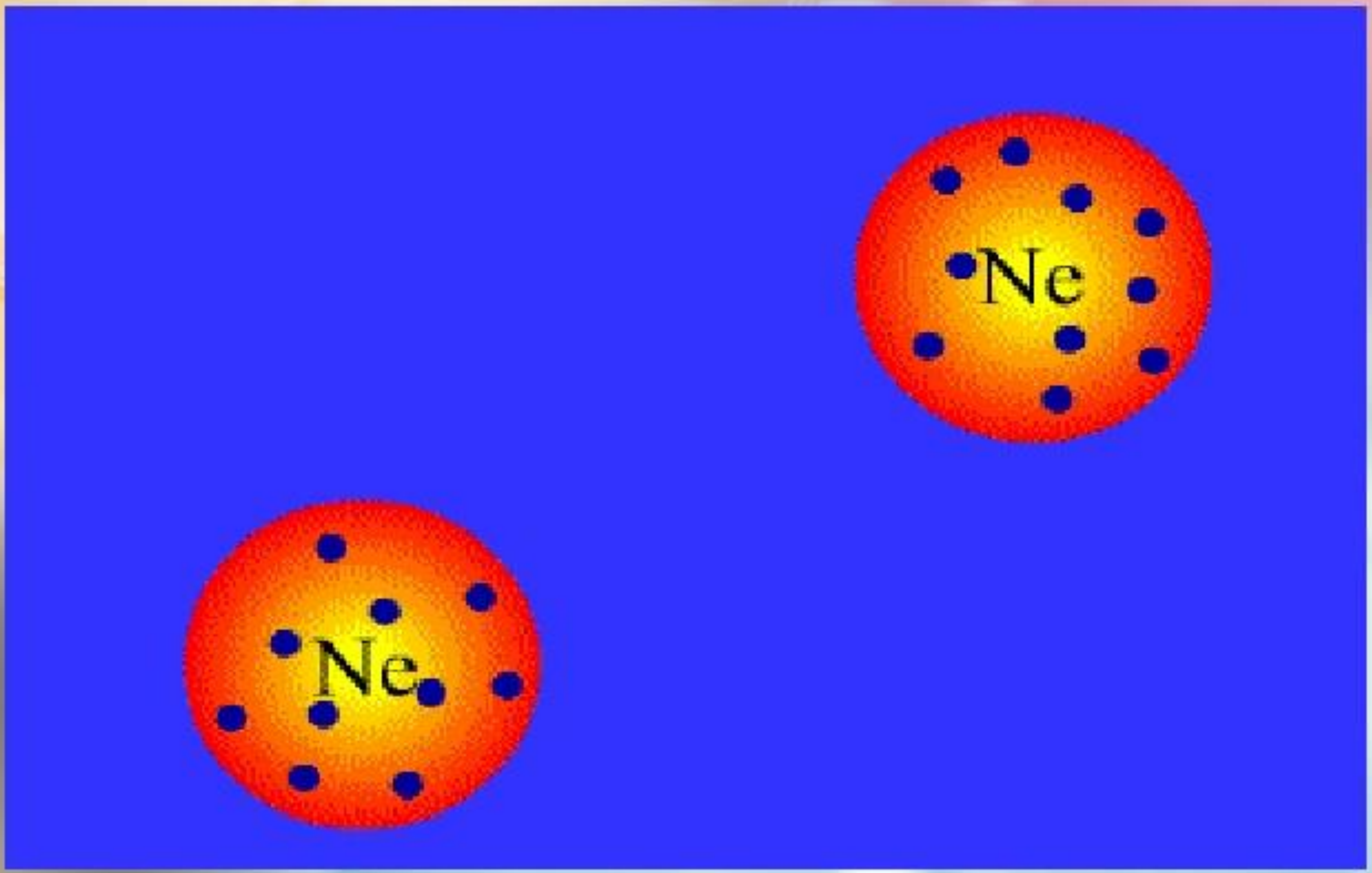
- van der Waals forces include
  - London forces
  - Dipole - dipole forces
  - Dipole - induced dipole forces
- Other intermolecular forces are
  - Ion - dipole interactions
  - Ion - induced dipole interactions
  - Hydrogen Bonding

LONDON FORCES (DISPERSION FORCES)

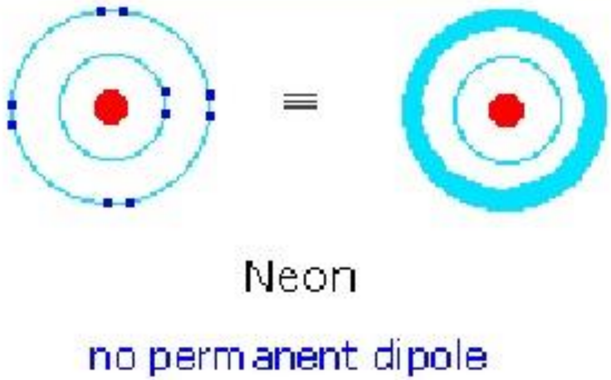
These arise from temporary variations in electron density in atoms and molecules. At any instant , the electron distribution may be unsymmetrical and hence produce an instantaneous dipole. This can cause an induced transient dipole in the neighbouring m
be



LONDON FORCES



- Dispersion forces are present between all molecules , whether polar or non-polar.
- Dispersion forces are stronger in molecules that are easily polarizable.
- Larger and heavier molecules exhibit stronger dispersion forces than smaller and lighter ones.

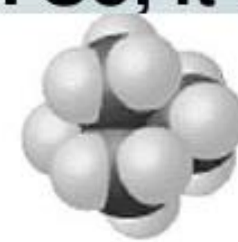


MAGNITUDE OF DISPERSION FORCES

- The cylindrical shape of n-pentane has greater surface area .Hence, n-
  on

n-Pentane
(bp = 309.4 K)

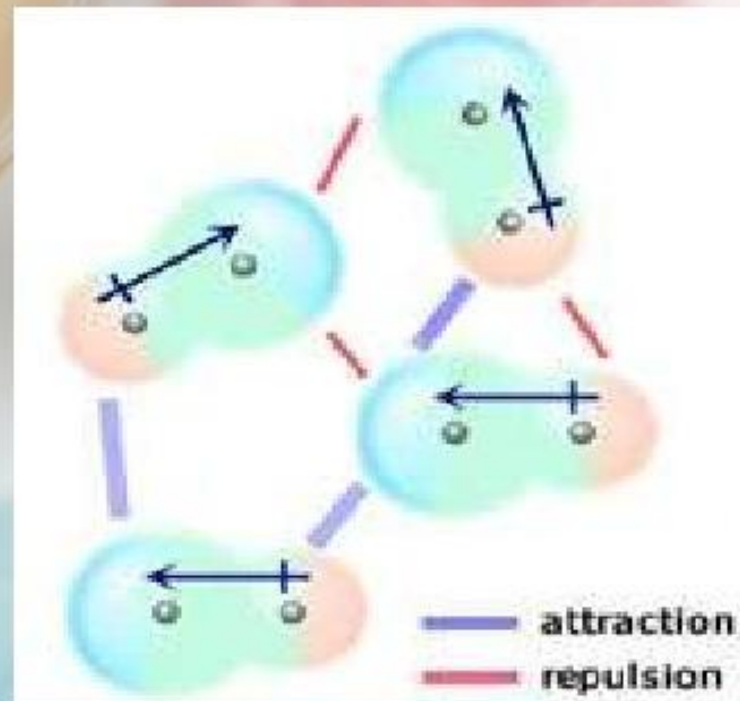
- The spherical shape of neo-pentane has lesser surface area. So, it has
  les
  att



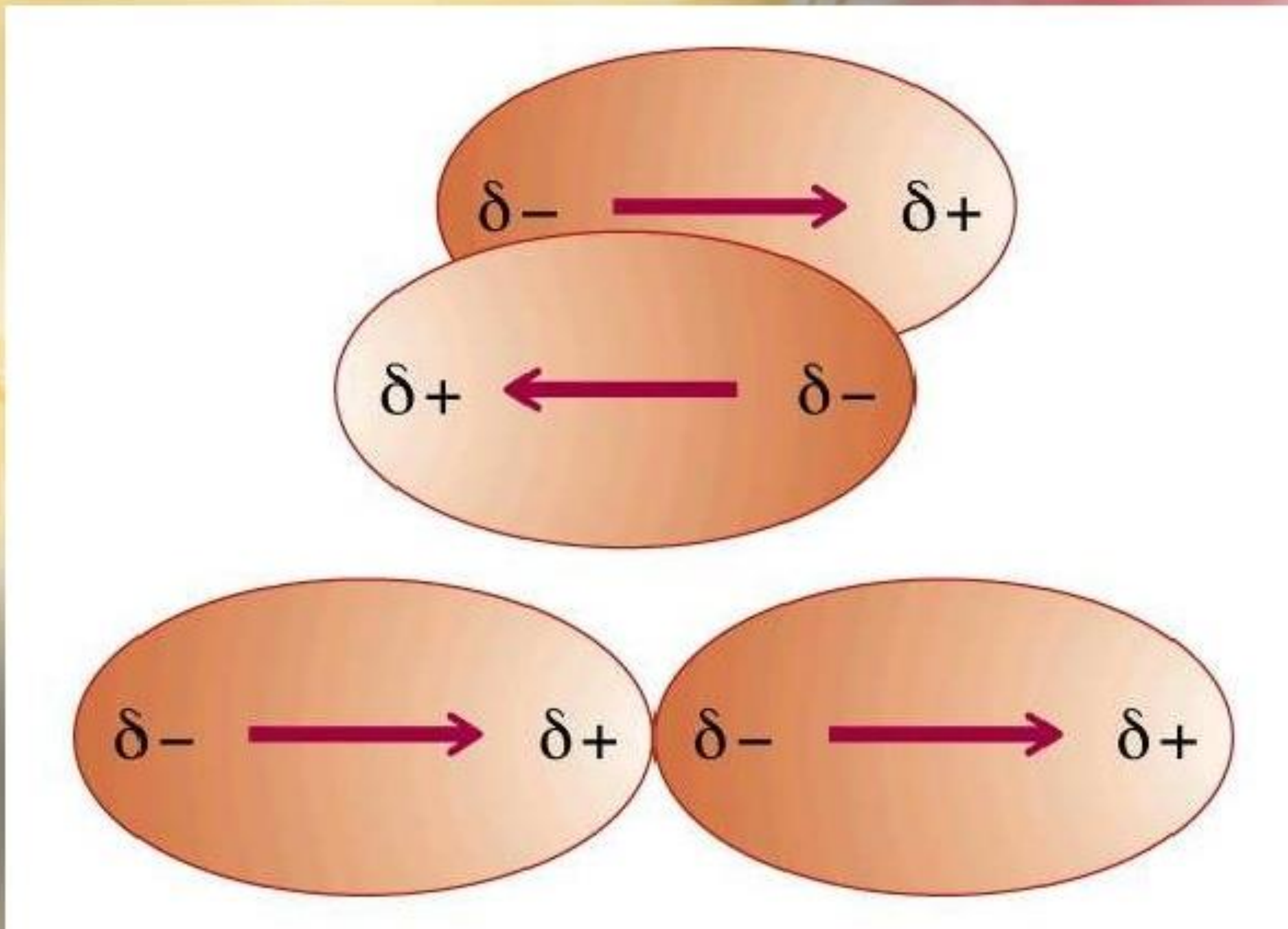
Neopentane
(bp = 282.7 K)

DIPOLE-DIPOLE FORCES(KEESOM FORCES)

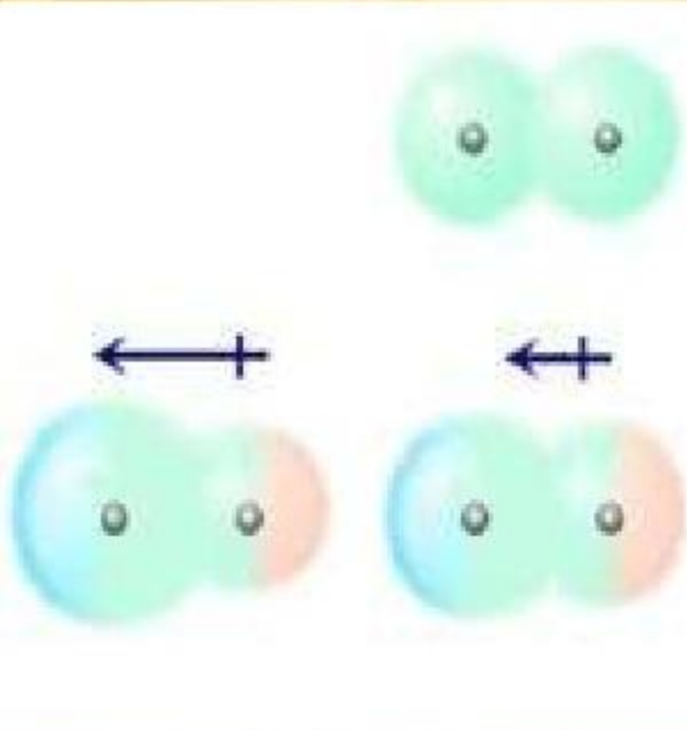
These forces arise due to interaction between oppositely charged ends of polar molecules. Greater the dipole moment of molecules , greater are the



FORCES



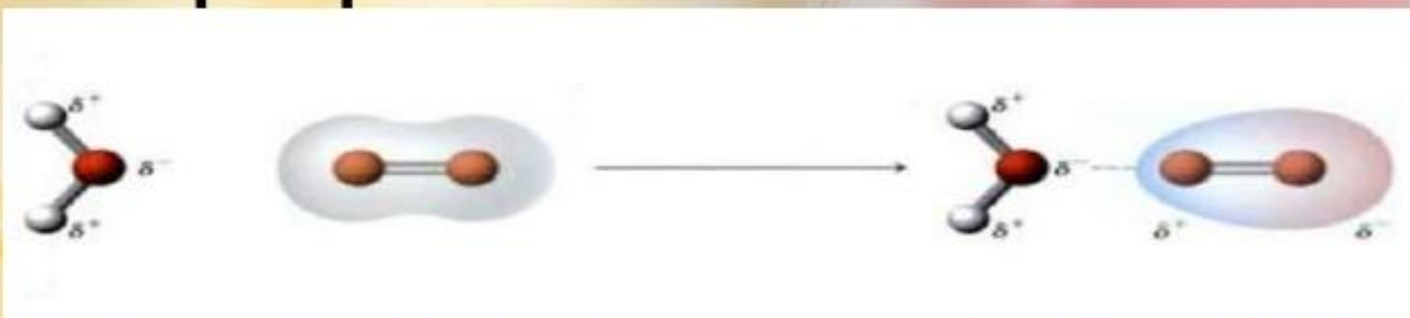
DIPOLE-INDUCED DIPOLE FORCES



These operate between polar molecules having permanent dipole and the molecules having no permanent dipole.The polar molecule induces a dipole in the neighbouring non-polar molecule.The interaction energy depends on

- Dipole moment of polar molecule

- Oxygen gas can dissolve in water because the permanent dipole in water can induce a dipole in oxygen



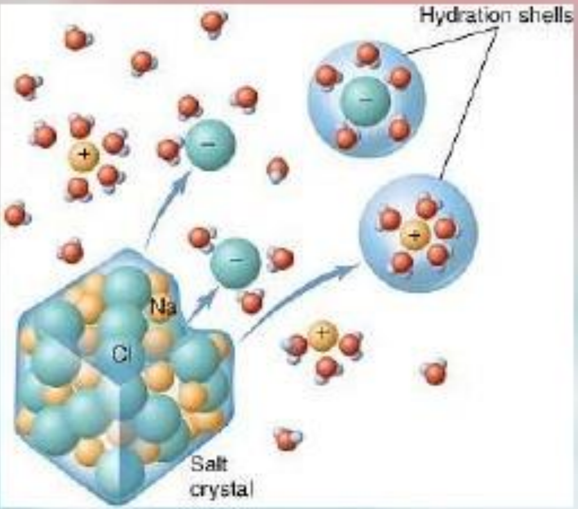
- **Non-polar iodine dissolves in polar ethanol due to dipole-induced dipole forces**

INTERACTIONS

These interactions depend on

- Charge and size of ion

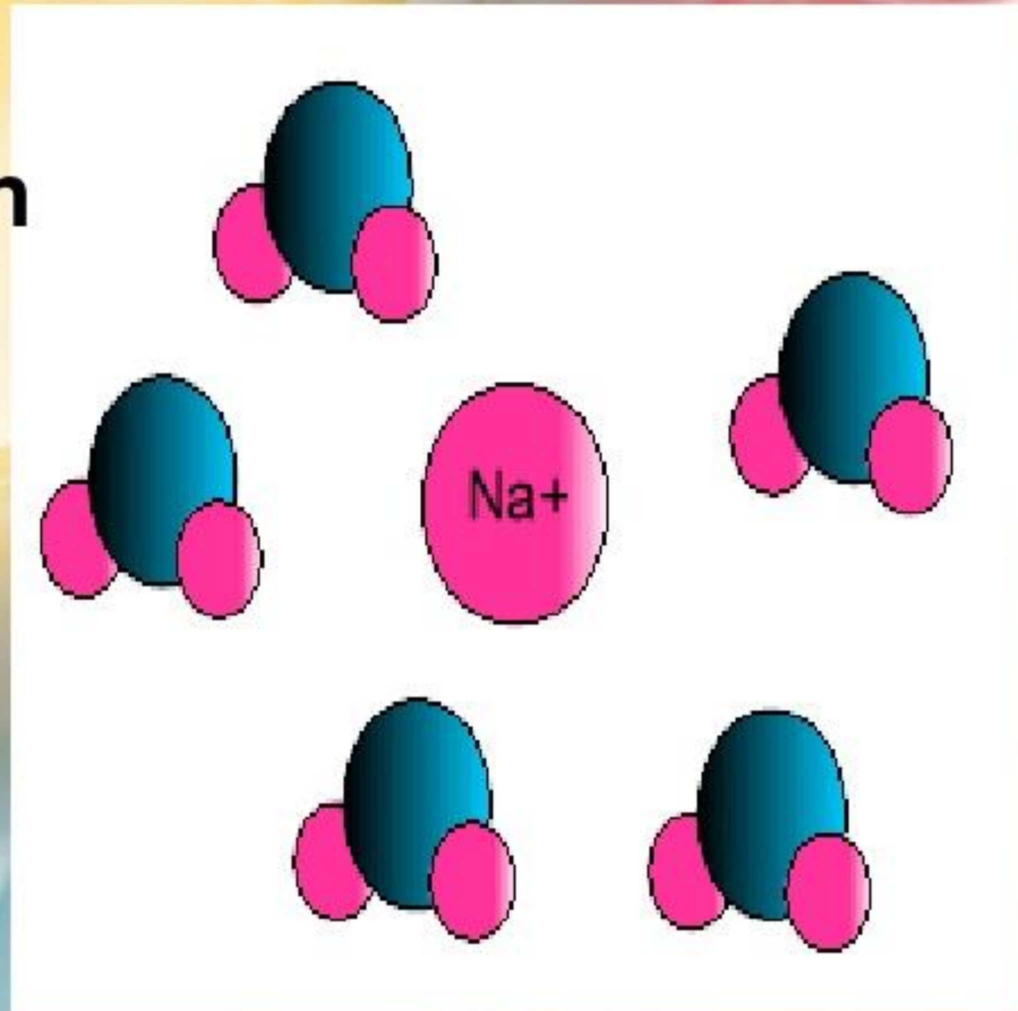
- Magnitude of dipole moment of polar molecule



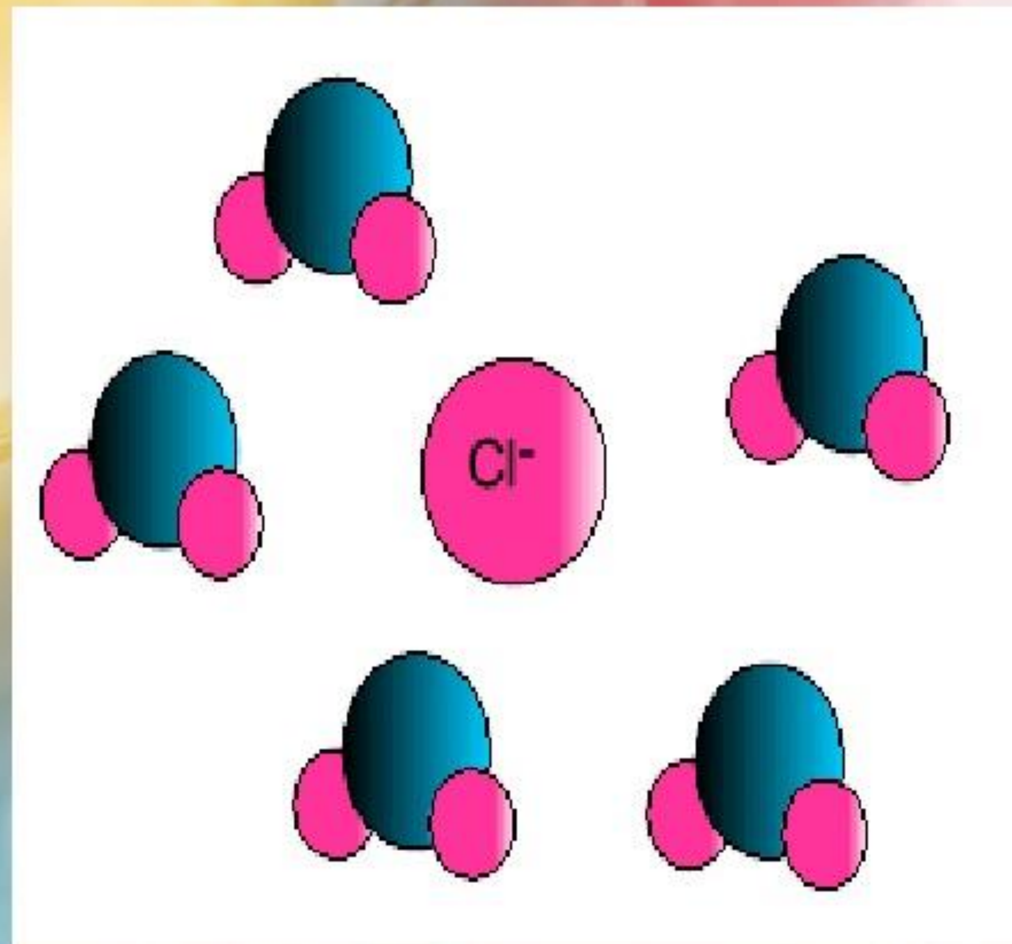
Hydration

Of

Sodium ions



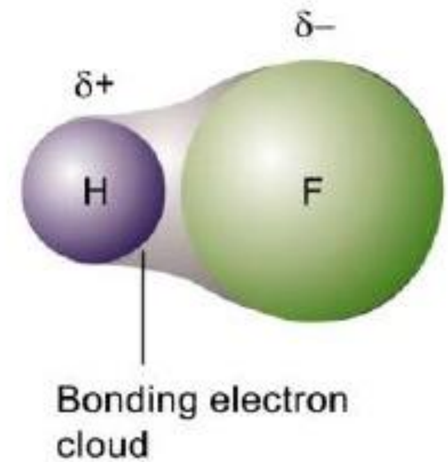
Hydration of Chloride Ions



HYDROGEN BONDING

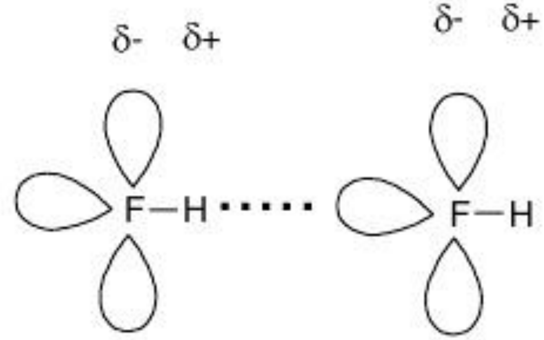
Hydrogen bonding

- When hydrogen is bonded to a highly electronegative atom (such as **nitrogen**, **oxygen**, **fluorine**), the bonding electron pair is drawn towards the **electronegative** atom



Hydrogen bonding

- Hydrogen has no **inner shell** electron and is very small in size, the positive charge density developed is **high**
- The nucleus of hydrogen atom is exposed to attraction by nearby electron cloud, a **lone pair electrons** on the electronegative atom



Hydrogen bonding

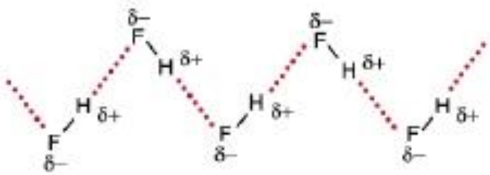
- Definition:
  - Electrostatic **attraction** between **hydrogen** atoms bonded to small, strongly **electronegative** atoms (N, O and F) and the **lone pair electrons** on these electronegative atoms

Hydrogen Bonding

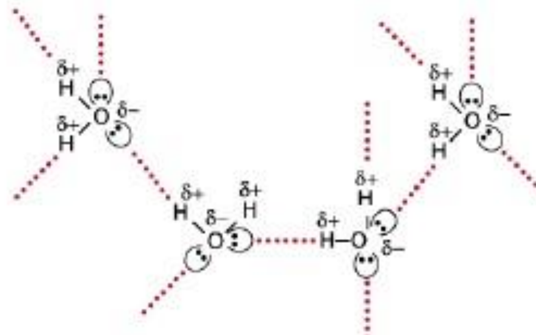
- Essential requirements for formation of hydrogen bond:
  - A hydrogen atom **must be directly bonded to a highly electronegative atom** (e.g. F, O and N)
  - **An unbonded pair of electrons (lone pair electrons)** is presented on the electronegative atom

Hydrogen bonding

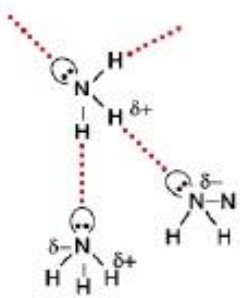
- Examples



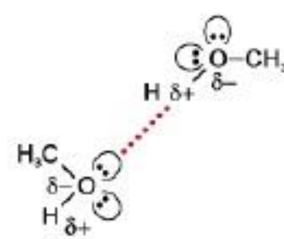
Hydrogen fluoride (HF)



Water (H_2O)



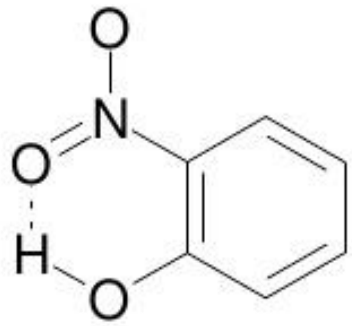
Ammonia (NH_3)



Methanol (CH_3OH)

Special Notes

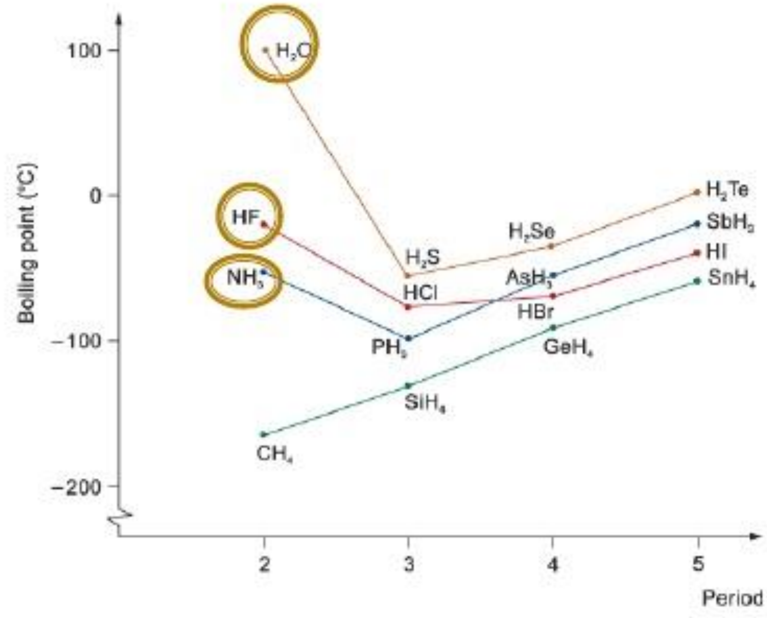
- Intermolecular hydrogen bond:
  - Hydrogen bond formed between two molecules
- Intramolecular hydrogen bond:
  - Hydrogen bond formed between two different atoms in the same molecule



- Intermolecular hydrogen bond is stronger than van der Waals' forces

Importance of Hydrogen Bonding in Physical Phenomena

- Anomalous Properties of the second period hydrides



Questions:

1. Why do H_2O, NH_3 and HF have abnormally high boiling point?

2. Explain why the order of boiling point is H_2O > HF > NH_3.

Q1: Why do NH_3, H_2O and HF have abnormally high boiling point?

- Explanation:
  - N, O and F are highly **electronegative** atoms
    - Formation of **intermolecular hydrogen bonds** in their hydrides.
  - Intermolecular hydrogen bonds are much **stronger** than van der Waals' forces
  - More energy is needed to break the hydrogen bonds in NH_3, H_2O and HF

Q2: Explain why the order of boiling point is $H_2O > HF > NH_3$

- Explanation

	No. of H atoms available for hydrogen bond formation	No. of lone pair electrons available	No. of hydrogen bonds can form per molecule
NH_3	3	1	1
H_2O	2	2	2
HF	1	3	1

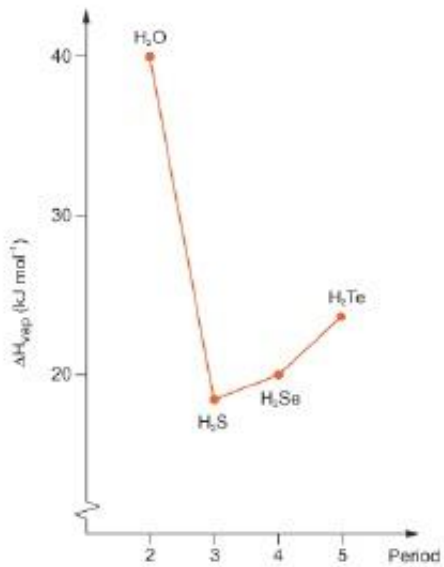
- H_2O can form **2** hydrogen bond per molecule
- NH_3 and HF can form only **1** hydrogen bond per molecule
- The boiling point of H_2O is higher than NH_3 and HF

Q2: Explain why the order of boiling point is $H_2O > HF > NH_3$

- Comparing N and F, **F** is more electronegative
- The intermolecular hydrogen bond formed between **HF** molecules is stronger than **NH_3**
- HF has a higher boiling point than NH_3

Importance of Hydrogen Bonding in Physical Phenomena

- Enthalpy of vaporization
  - Energy required to vaporized **one mole** of liquid

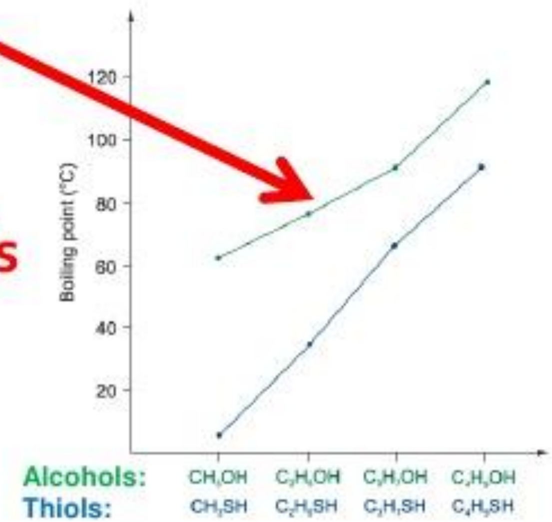


Question:
Explain why H_2O has the largest enthalpy of vaporization than H_2S, H_2Se and H_2Te.

- Explanation:
  - In H_2O, there are **intermolecular hydrogen bonds** between molecules
  - However, in H_2S, H_2Se and H_2Te, the interaction between molecules is **dipole-dipole interactions**
  - The strength of hydrogen bond is **stronger** than **dipole-dipole interactions**
  - To break the hydrogen bonds, **more** energy is required
  - H_2O has the largest enthalpy change of vaporization

Importance of Hydrogen Bonding in Physical Phenomena

- Boiling Point and Solubility of Alcohols
  - The boiling points of alcohols are higher than the thiols because there are **intermolecular hydrogen bonds** between alcohol molecules, but only **dipole-dipole interactions** between molecules of thiols

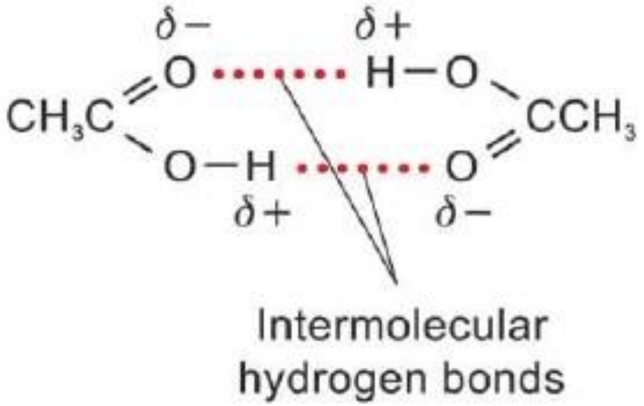


- Alcohols of low molecular masses are **soluble** in water because they can form **intermolecular hydrogen bonds** with water molecules

Importance of Hydrogen Bonding in Physical Phenomena

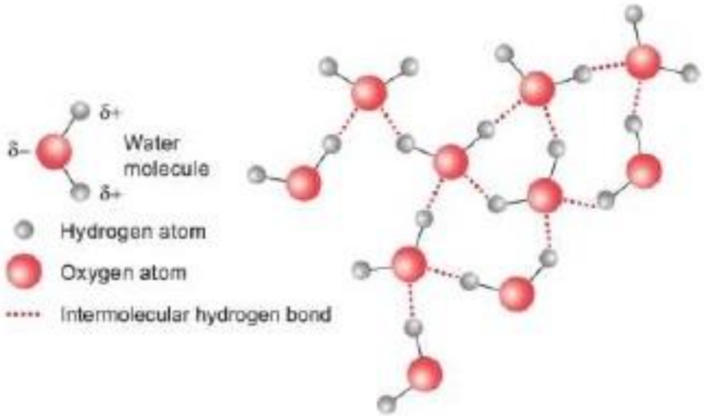
- Dimerization of Carboxylic Acid
  - It happens when carboxylic acids are dissolved in **non-polar** solvents or in **vapour** state

  E.g. Ethanoic acid (CH_3COOH)
    - Ethanoic acid molecules form **dimers** through the formation of **intermolecular hydrogen bonds**



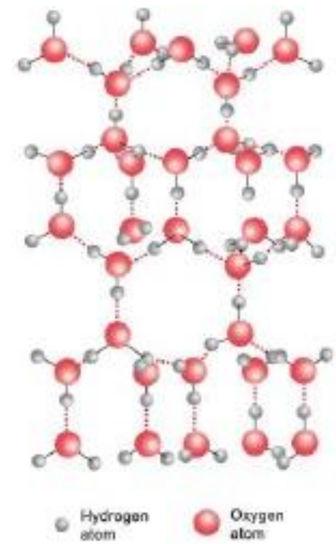
Importance of Hydrogen Bonding in Biochemistry

- Hydrogen bonding in ice and water
  - **Intermolecular hydrogen bonds** exist in both water and ice
  - **In water**, molecules are in constant motion. **Intermolecular hydrogen bonds** are formed and broken continuously.
  - The arrangement of molecules is **random**



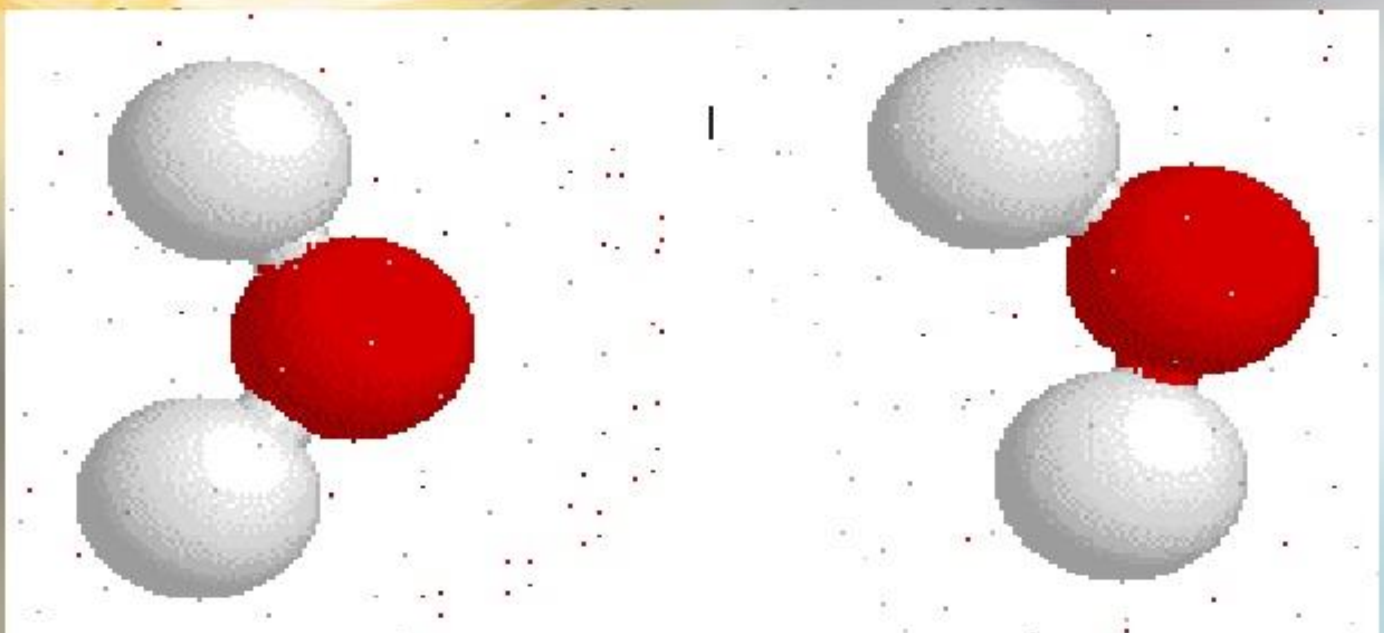
- **In ice**, the molecular motion is **restricted**. The molecules are arranged in a way to form the **maximum** number of hydrogen bonds

- Each H_2O molecule is **tetrahedrally** bonded to 4 other H_2O molecules by hydrogen bonds
  - Creates an **open structure** and gives rise to a **lower** density of ice than water
  - Insulates the water below and prevents complet solidification ➔ keep fish survive in polar regions



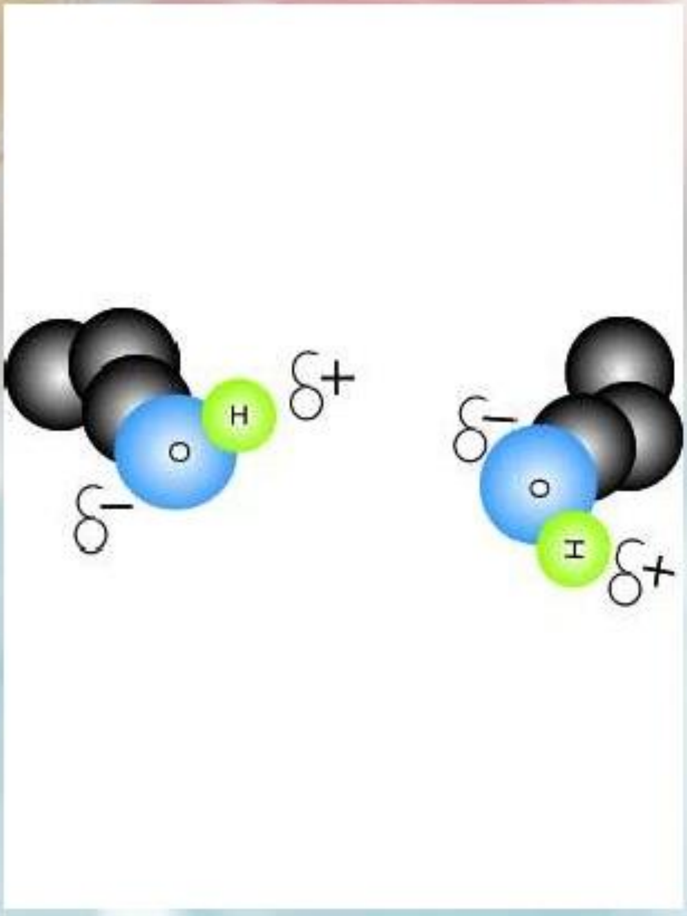
HYDROGEN BONDING

- It is an electrostatic force of attraction that exists between covalently bonded hydrogen atom of one molecule and the electronegative atom of another molecule.



CONDITIONS FOR HYDROGEN BOND FORMATION

- A hydrogen atom attached to a relatively electronegative atom is a hydrogen bond donor. This electronegative atom is usually fluorine, oxygen, or nitrogen.
- An electronegative atom such as fluorine, oxygen, or nitrogen is a



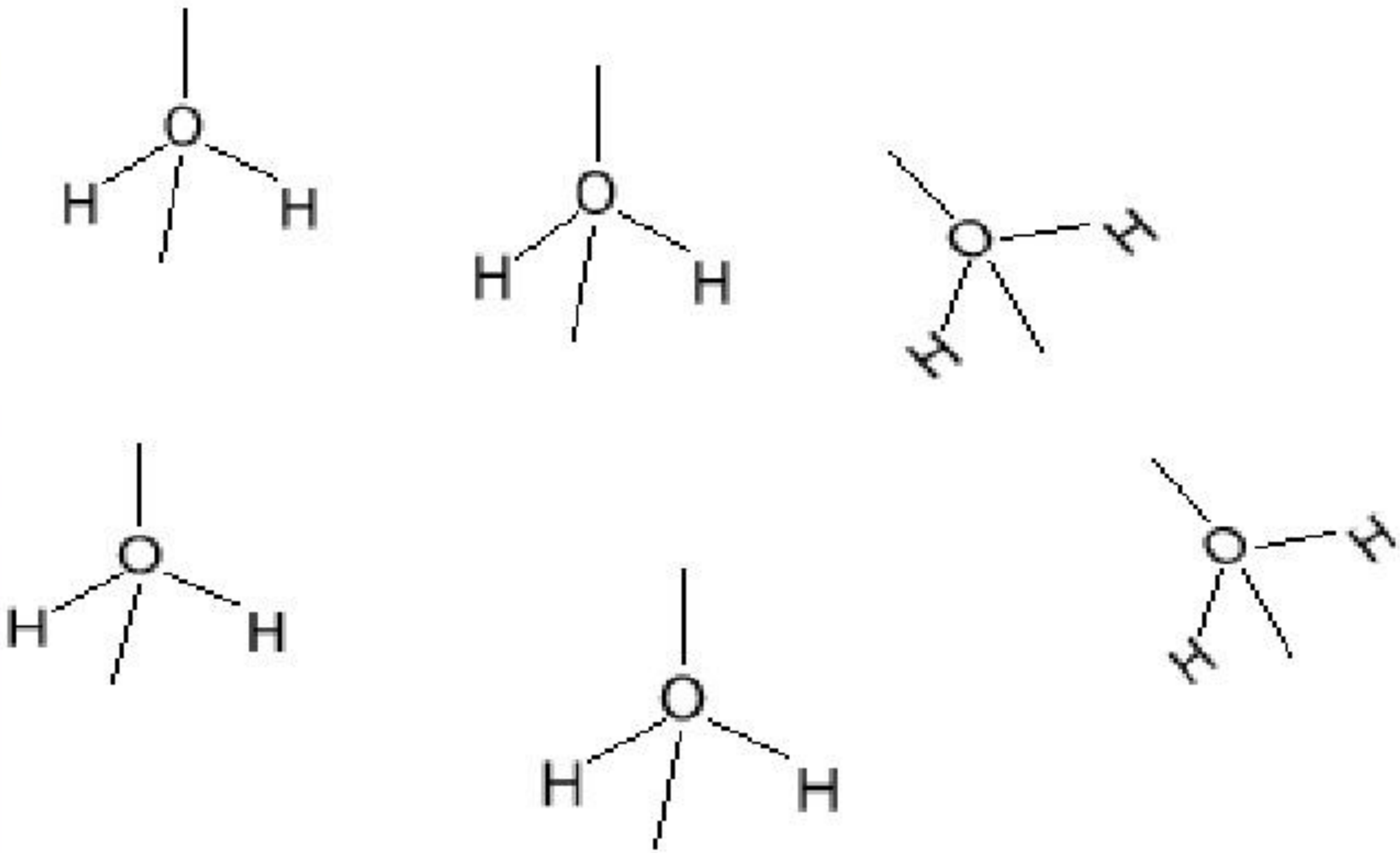
INTERMOLECULAR HYDROGEN BOND

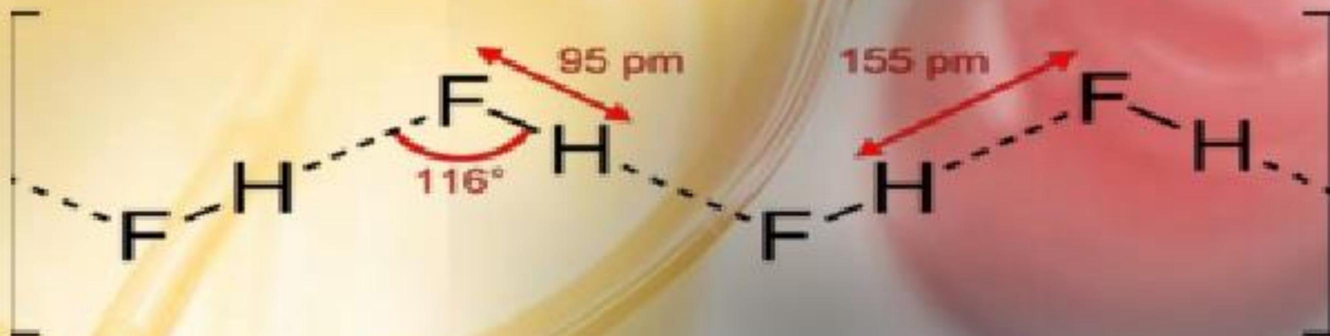
- It is formed between two different molecules of the same or different substances, eg.
- Hydrogen bond between molecules of ammonia
- Hydrogen bond between molecules of water and alcohol

INTRAMOLECULAR HYDROGEN BOND

- It is formed between the hydrogen atom and a highly electronegative atom present in different bonds within the same molecule, eg.
- o-salicylaldehyde
- o- nitrobenzoic acid

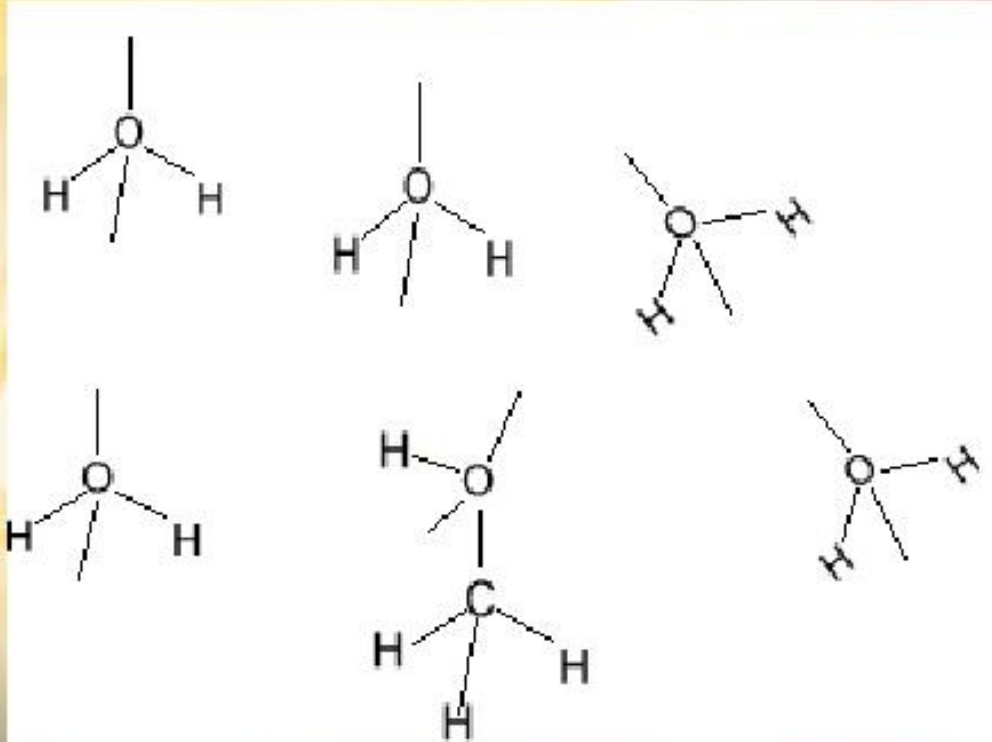
ASSOCIATION OF WATER MOLECULES





HYDROGEN FLUORIDE

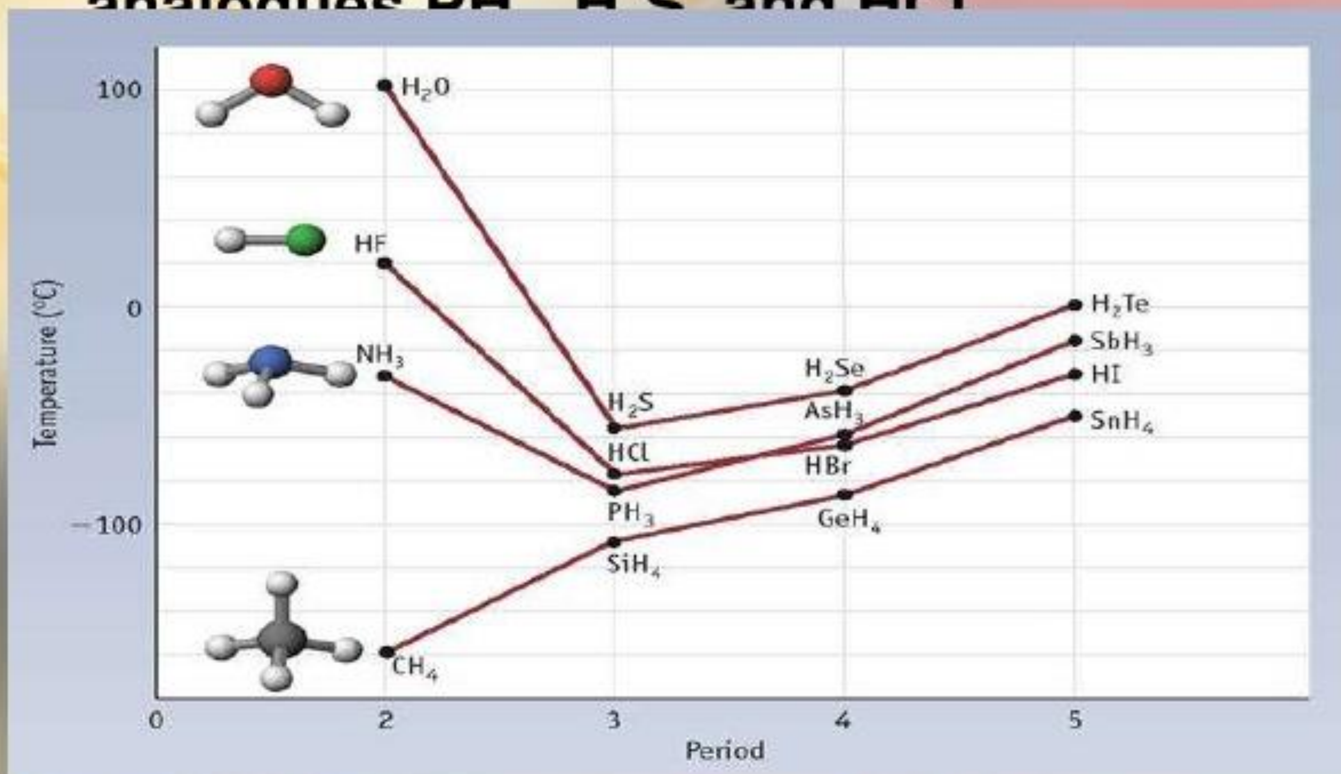
IN THE SOLID STATE , HYDROGEN FLUORIDE CONSISTS OF LONG ZIG-ZAG CHAINS OF MOLECULES ASSOCIATED TOGETHER BY HYDROGEN BONDING. HENCE, HYDROGEN FLUORIDE MOLECULE CAN ALSO BE REPRESENTED AS $(HF)_N$. HOWEVER, IN GASEOUS OR LIQUID STATE , THE CHAIN BECOMES LINEAR.



ALCOHOLS HAVE HIGH MISCIBILITY IN WATER DUE TO FORMATION OF INTERMOLECULAR HYDROGEN BONDS BETWEEN WATER AND ALCOHOL MOLECULES.

HIGHER MELTING AND BOILING POINTS

- Dramatically higher boiling points of NH_3, H_2O, and HF compared to the heavier analogues PH_3, H_2S, and HCl

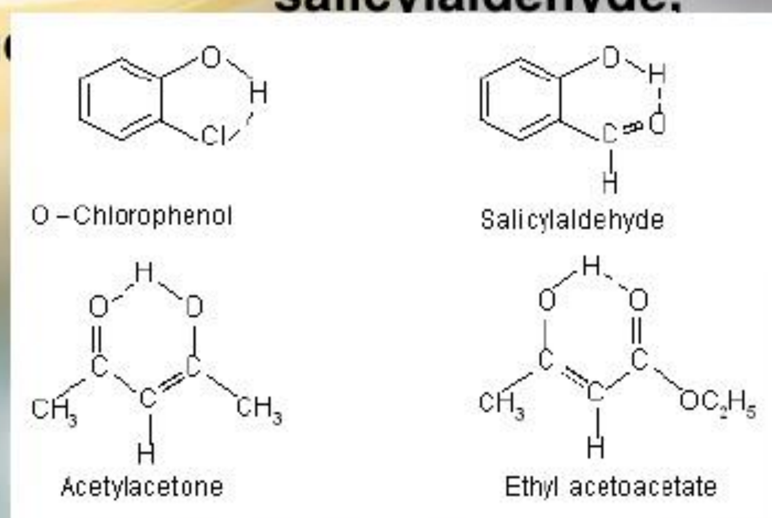


HYDROGEN BONDING CAUSES

- Higher boiling point in ethanol as compared to diethyl ether
- Higher viscosity of sulphuric acid and glycerol
- Water is a liquid, whereas H_2S is a gas
- Dimerisa… molecules

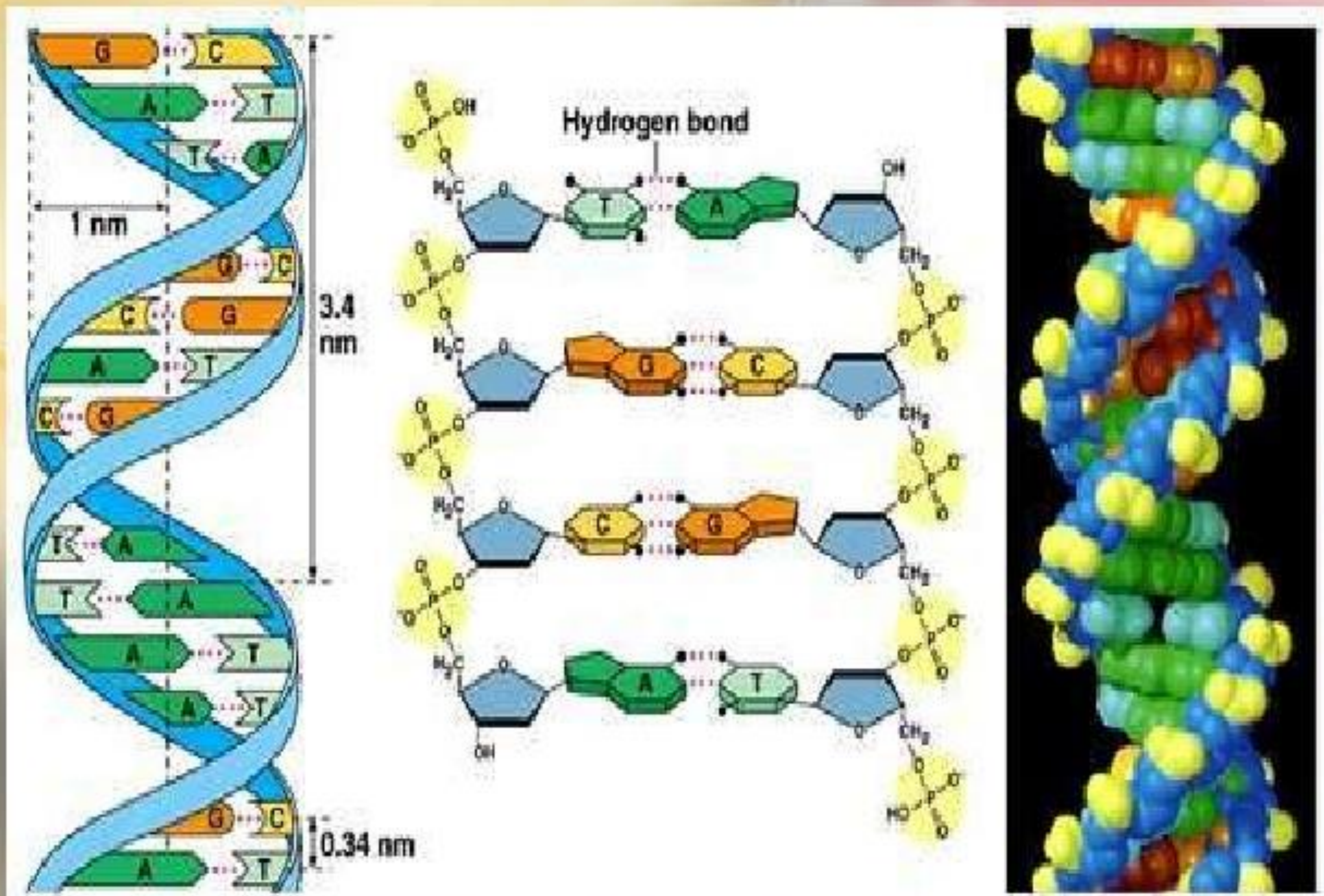
INTRAMOLECULAR HYDROGEN BONDING

When hydrogen bonding exists within the molecule it is called intramolecular hydrogen bonding. In such type of hydrogen bonding two groups of the same molecule link through hydrogen bond, forming a stable five or six membered ring structure *e.g.,* salicylaldehyde, *o*-chloroph acetate etc.

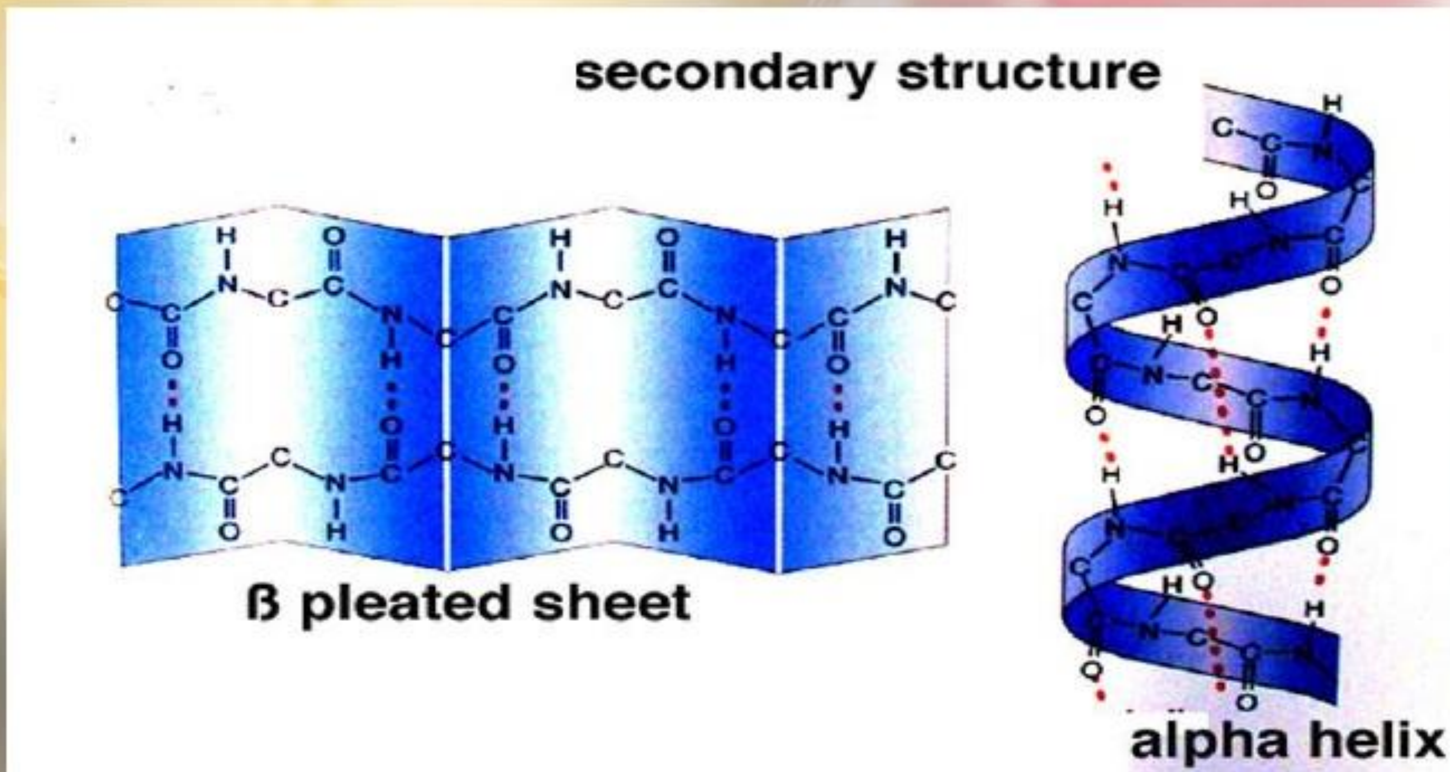


This intramolecular hydrogen bonding was first called chelation (after the Greek word "Chela" meaning, claw) because in the same molecule the formation of a ring hydrogen bonding is a pincer like action resembling the closing of a Crab's claw. Some more examples of intramolecular hydrogen bonding are :

HYDROGEN BONDING IN DNA



SECONDARY STRUCTURE OF PROTEINS-HYDROGEN BONDING

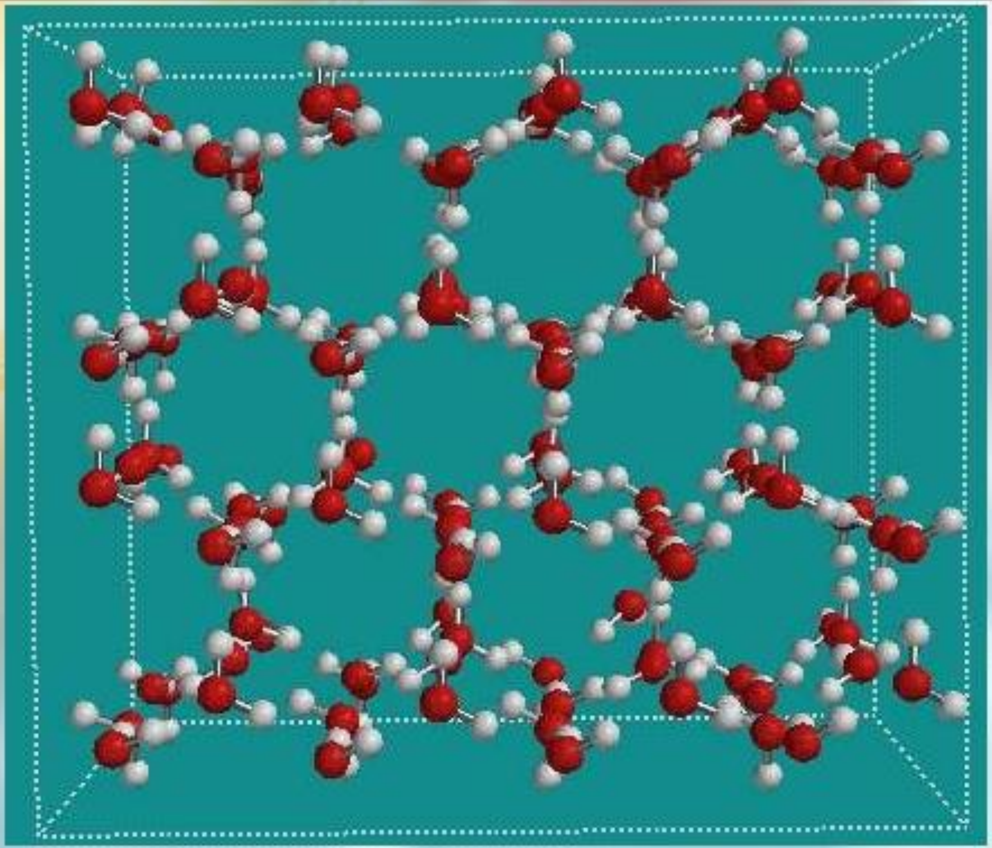


CAGE LIKE STRUCT URE OF

WATER MOLECU LES

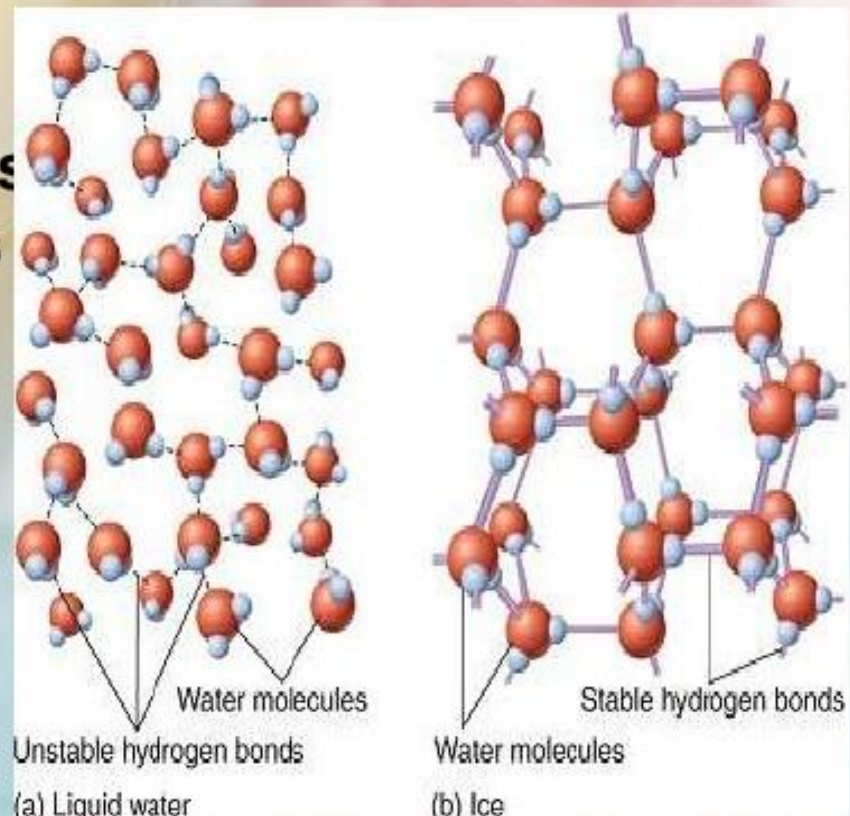
IN ICE

CAUSES LOWER DENSITY OF ICE

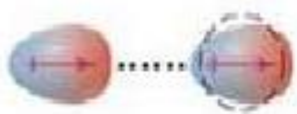


UNIQUE NATURE OF WATER

- When ice melts, this cage like structure collapses and the molecules move closer.
- Hence, for the same mass of water, volume decreases and density increases. Hence ice has lower density than water at 273K

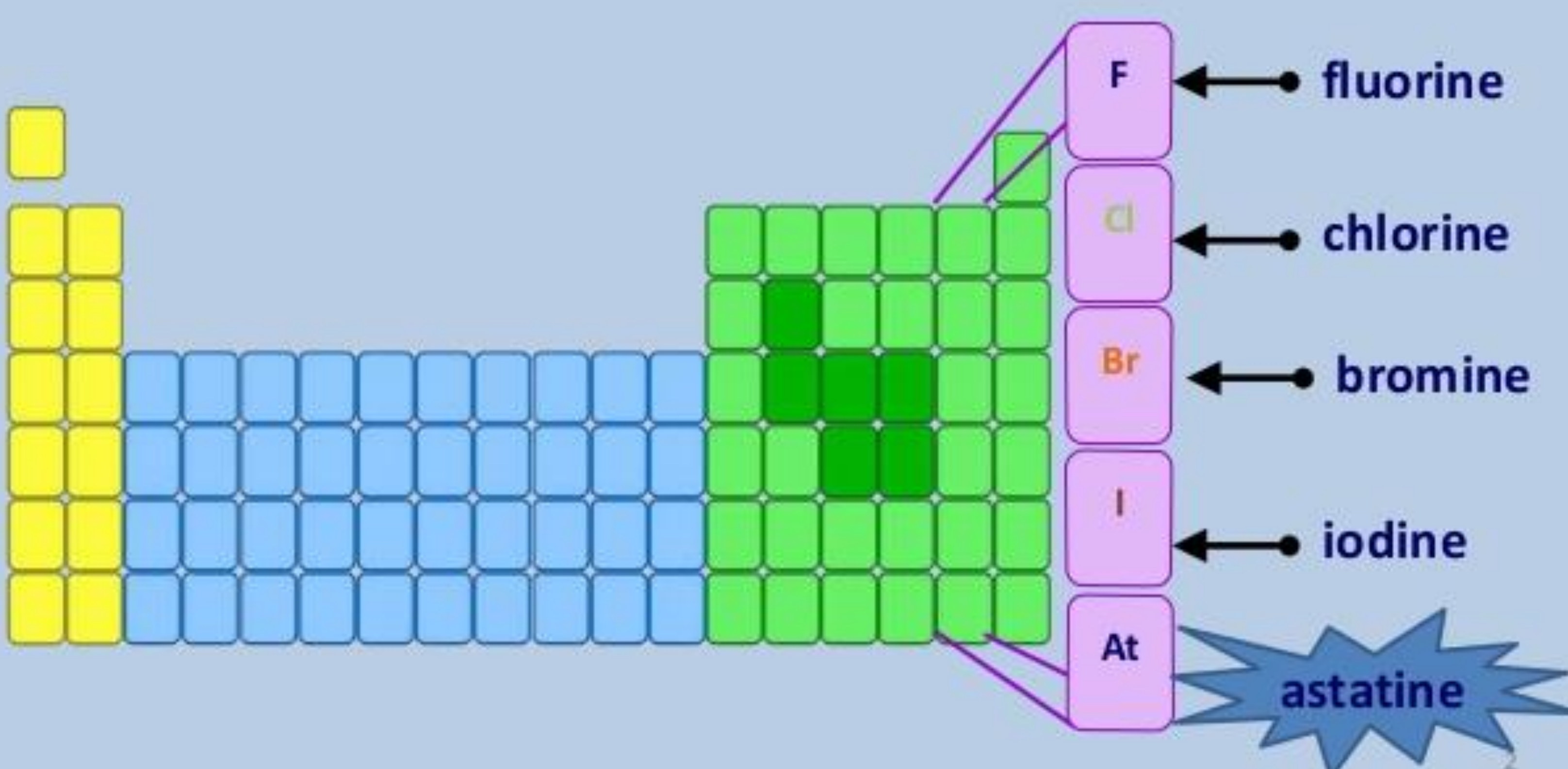


COMPARISON OF INTERMOLECULAR FORCES

Force	Model	Basis of Attraction	Energy (kJ/mol)	Example
Ion-dipole		Ion charge–dipole charge	40–600	$Na^{+}\cdots O(H)_2$
H bond	δ- δ+ δ- —A—H·······:B—	Polar bond to H–dipole charge (high EN of N, O, F)	10–40	:Ö—H···:Ö—H (each O bonded to H)
Dipole-dipole		Dipole charges	5–25	I—Cl···I—Cl
Ion–induced dipole		Ion charge–polarizable e^- cloud	3–15	$Fe^{2+}\cdots O_2$
Dipole–induced dipole		Dipole charge–polarizable e^- cloud	2–10	H—Cl···Cl—Cl
Dispersion (London)		Polarizable e^- clouds	0.05–40	F—F···F—F

INTER HALOGEN COMPOUNDS

The elements in group 17 of the periodic table, **on the right**, are called the **HALOGENS**.



Why the name ..... HALOGEN??

* Halogen-metal compounds are salts occurring in sea water

(e.g. NaCl; sodium chloride),

* *halos* = sea salts; *genes*=born.

Electron configuration of halogens

Fluorine	$1s^2\ 2s^2\ 2p^5$
Chlorine	$[Ne]3s^2\ 3p^5$
Bromine	$[Ar]3d^{10}\ 4s^2\ 4p^5$
Iodine	$[Kr]4d^{10}\ 5s^2\ 5p^5$
Astatine	$[Xe]4f^{14}\ 5d^{10}\ 6s^2\ 6p^5$

Interhalogen Compounds

- These covalent compounds are formed when two different halogens react .
- These are formed due to the electronegativity difference among the halogens.

Classification:

- Interhalogens have the general formula Ax_n where n=1, 3, 5 &7.

Type AX eg: ClF, BrF, BrCl, ICl, Ibr

Type $AX_{3:}$ eg : $ClF_{3,}$ $BrF_{3,}$ ICl_3

Type AX_5 eg: BrF_5 ClF_5 Type AX_7 eg: IF_7

When representing the compound, the less electro negativity element has to be written first.

Iodine mono chloride ICl

is formed by passing chlorine over solid iodine at temperature below 0 ^{0}C.

$$I_2 + Cl_2 \longrightarrow 2ICl$$

- It is a red-brown chemical compound
- melts near room temperature.
- Because of the difference in the electronegativity of iodine and chlorine, **ICl** is highly polar ; I^+Cl^-

* In organic synthesis, estimation of iodine No. of oils and as a source of I^+.

1. Iodine monochloride, ICl:

Pretaration:

- It is formed by mixing I_2 and Cl_2 in equal amounts.

  $I_2 + Cl_2$ (equal amounts) $\rightarrow$ 2 ICl
- By heating I_2 with $KClO_3$

  $KClO_3 + I_2 \rightarrow KIO_3 + ICl$
- By heating ICl_3 at 68^0 C

Properties:

i. It is dark liquid with b. p. 97.4^0 C.

ii. ICl exists in two solid form as given below:

- Solid form: This form melts at 27.2^0 C and is obtained as needle like crystals on cooling the liquid rapidly.
- Metastable form: This form melts at 14^0 C and is obtained as a black solid on cooling the liquid slowly at 100^0 C.

iii. Hydrolysis: ICl dissolve in water and gets hydrolyzed as:

$$ICl + H_2O \rightleftharpoons IOH + HCl$$

$$3\ ICl + 3\ H_2O \rightleftharpoons HIO_3 + 3\ HCl + 3\ HI$$

$$5\ ICl + 3\ H_2O \rightleftharpoons HIO_3 + 5\ HCl + 2\ I_2$$

The hydrolysis can be prevented adding HCl

iv. ICl decomposed by excess of KOH as:

$3\ ICl + 6\ KOH \rightarrow 3\ KCl + 2\ KI + KIO_3 + 3\ H_2O$

v. In liquid state ICl undergoes auto-ionization as:

$2\ ICl \rightleftharpoons I^+$ (Solvent cation) $+ ICl_2^-$ (Solvent anion)

Metal halide give I^+ ions in liquid ICl hence it behaves as Lewis acid.

$AlCl_3 + ICl \rightleftharpoons I^+ + AlCl_4^-$; $TiCl_4 + ICl \rightleftharpoons I^+ + TiCl_5^-$

$MCl_5 + ICl \rightleftharpoons I^+ + MCl_6^-$; (M=Sb, Rb)

Alkali metal chloride gives ICl_2^- ion in liquid ICl and therefore acts as bases.

$MCl + ICl \rightleftharpoons M^+ + ICl_2^-$, (M=K, Rb)

vi. Bromide like KBr, $[(CH_3)_4N]Br$ etc. , reacts with ICl, polyhalide ion is produced.

$ICl + KBr \rightleftharpoons K^+[ClBrI]^-$

$ICl + [(CH_3)_4N]Br \rightleftharpoons [(CH_3)_4N]^+[ClBrI]^-$

vii. Excess of Cl_2 convert ICl into ICl_3

Cl_2 (excess) $+ ICl \rightarrow ICl_3$

viii. When molten ICl is electrolyzed, a mixture of I_2 and Cl_2 is liberated at anode and I_2 is liberated at cathode. This suggests that ICl ionizes as:

$2\ ICl \rightleftharpoons I^+$ (Solvent cation) $+ ICl_2^-$ (Solvent anion)

Reaction at cathode: $2I^+ + 2\ e^- \rightarrow I_2$ (reduction)

Reaction at anode : $2\ ICl_2^- \rightarrow I_2 + 2\ Cl_2 + 2\ e^-$ (oxidation)

Uses:

- A solution of ICl is used as catalyst in the oxidation of As(III) oxide by ceric sulphate and for preparation of polyhalides.
- The solution of ICl in glacial acetic acid is used for determining the iodine value of oil by Wiz's method.

Structure of ICl- Its linear in shape

Bromine trifluoride... BrF_3

- It is obtained by mixing bromine vapor and fluorine in a stream of nitrogen at 20°C.

$$Br_2 + 3F_2 \longrightarrow 2BrF_3$$

It is a straw-colored liquid with a pungent odor.

It is a powerful fluorinating agent

It is used to produce uranium hexafluoride (UF_6) in the processing and reprocessing of nuclear fuel.

Preparation:

- It is obtained by mixing Br_2 vapour with F_2 in presence of N_2

$$Br_2(g) + 3F_2 \xrightarrow{N_2} 2BrF_3$$

- By action of ClF_3 on Br_2 at 10^0 C.

$$ClF_3 + Br_2 \xrightarrow{10^0 C} BrF_3 + BrCl$$

Properties:

i. It is fuming liquid with b. p.125.8^0 C, is very reactive and non corrosive liquid.

ii. It reacts with Br_2 to form BrF : $BrF_3 + Br_2 \rightarrow 3BrF$

iii. *Auto-ionisation:* $2BrF_3 \rightleftharpoons BrF_2^+ + BrF_4^-$

The substance making available BrF_2^+, due to above mode of auto-ionisation; acts as acids.

Examples: $AuF_3 + BrF_3 \rightleftharpoons BrF_2^+ + AuF_4^-$

MF_4 (M=Ge, Sn, Ti) + $2BrF_2^+ \rightleftharpoons 2BrF_2^+ + MF_6^{2-}$

The substance making available BrF_4^- acts as bases.

Examples: MF (M=Li, K, Ag) + $BrF_3 \rightleftharpoons M^+ + BrF_4^-$

iv. *Neutralisation reactions:* These reactions are those in which a compound containing BrF_2^+ ion reacts with that having BrF_4^- ion and salt and solvent are formed.

Acid Base Salt Solvent

v. BrF_3 is useful *fluorination agent* as follow.

$6\,M + 2\,BrF_3 \rightarrow 6\,MF + Br_2$; $3\,MCl + BrF_3 \rightarrow 3\,MF + BrCl_3$

$WO_3 + 4\,BrF_3 \rightarrow 2\,WF_6 + 2\,Br_2 + 3\,O_2$

$6\,CuO + 4\,BrF_3 \rightarrow 6\,CuF_2 + 3\,Br_2 + 3\,O_2$

vi. *Redox reactions and complex formation:*

$3\,K + 4\,BrF_3 \rightarrow 3\,K^+[BrF_4]^- + \frac{1}{2}\,Br_2$

zxc

vii. *Solvolysis reaction:* In these reaction the concentration of $[BrF_2]^+$ and $[BrF_4]^-$ is increased.

$4\,BrF_3(l) + K_2[TiF_6] \rightleftharpoons [BrF_2]_2[TiF_6] + 2\,K[BrF_4]$

$2\,BrF_3(l) + K[PF_6] \rightleftharpoons [BrF_2][PF_6] + K\,[BrF_4]$

viii. *Formation of adducts:*

$XeF_2 + MF_5 \rightarrow XeF_2.MF_5$ (In BrF_3 solution)

BrF_3 hydrolysed by water with formation of bromine oxy-fluoride.

$BrF_3 + H_2O \rightarrow BrOF_3 + H_2$

Uses:

- For preparation of polyhalides.
- For preparation of complex compounds by acid-base reaction, are difficult to prepare by other methods.
- As fluorinating agent.
- For preparation of fluorocomplexes of many metals like Au, Ge, As etc.

Molecule : BrF_3

$$x = \frac{1}{2}[7+3]$$

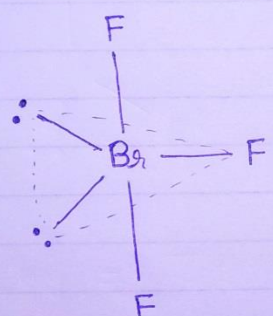
$$x = \frac{10}{2}$$

$$x = 5$$

$\therefore$ Hybridisation = sp^3d ⇒ Trigonal bipyramidal

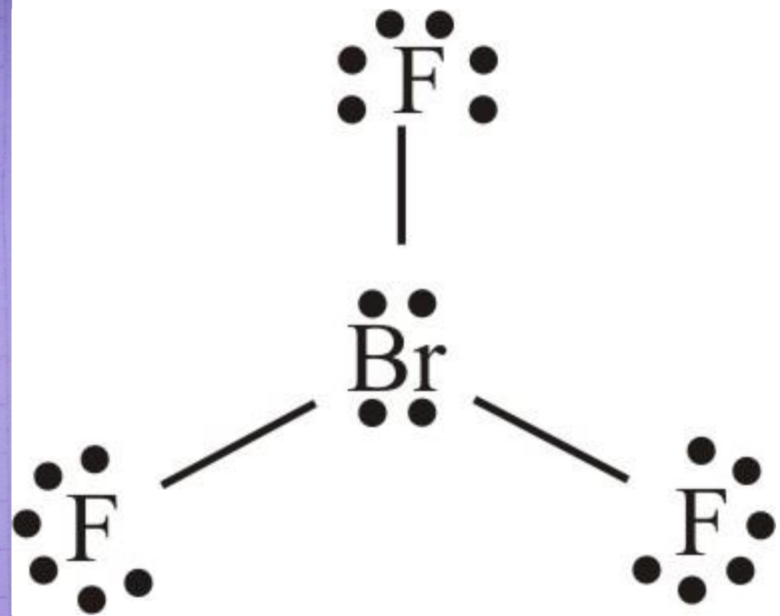
l.p. = 5 − 3 = 2

b.p. = 3



Geometry = Trigonal bipyramidal

Shape = T-shaped



2. Iodine pentafluoride, IF_5:

Preparation:

- It is formed by direct combination of I_2 and F_2 (excess):

  $I_2 + 5 F_2 \text{ (excess)} \rightarrow 2 IF_5$

- By the action of F_2 on I_2O_5: $10 F_2 + 2 I_2O_5 \rightarrow 4 IF_5 + 5 O_2$
- By heating I_2 with AgF: $I_2 + 10 AgF \rightarrow 2 IF_5 + 10 Ag$

Properties:

i. It is courless liquid with m. p. = 9.6^0 C.

ii. It is good conductor of electricity as it ionizes as:

$$2 IF_5 \rightleftharpoons IF_4^+ + IF_6^-$$

Thus, the substances that give IF_4^+ ions in liquid IF_5, acts as acid and those which produce IF_6^- ions behave as bases in this solvent. The following acid-base reaction takes place in liquid IF_5.

$$[IF_4]^+[SbF_6]^- \text{(Acid)} + K[IF_6]^- \text{ (Base)} \rightleftharpoons K^+[SbF_6]^- \text{ (Salt)} + 2 IF_5 \text{ (Solvent)}$$

iii. IF_5 reacts with I_2O_5 to form iodine oxy-fluoride, IOF_5 which is solid and decomposes on heating at 110^0 C.

$$3 IF_5 + I_2O_5 \longrightarrow 5 IOF_3; \quad 2 IOF_3 \xrightarrow{\Delta} IF_5 + IO_2F$$

i. Hydrolysis of IF_5 gives halogen acid and oxy-halic acid.

$$IF_5 + 3 H_2O \rightarrow 5 HF + HIO_3$$

v. With F_2 at 250^0-300^0 C temperature gives IF_7.

$$IF_5 + F_2 \xrightarrow{250^0\text{-}300^0 C} IF_7$$

<u>USES</u>

It is used as an ionising solvent.

IF_5

Hybridisation of Central atom $= \frac{1}{2}[V + M - C + A]$

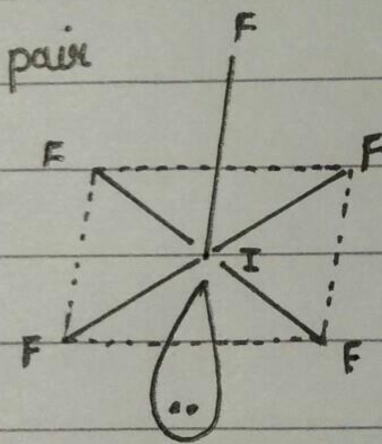
$H = \frac{1}{2}[7 + 5 - 0 + 0] = \frac{12}{2}$

$H = 6 = sp^3d^2$

$\therefore$ $7e^-$ → $5e^-$ (5 bond pair)

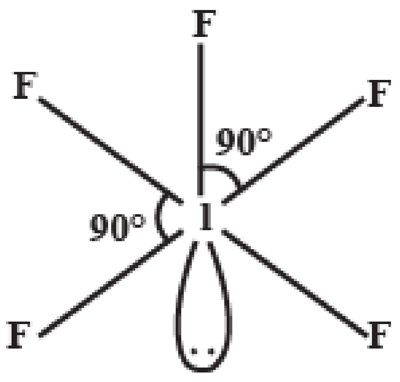
$7e^-$ → $2e^- = 1$ lone pair

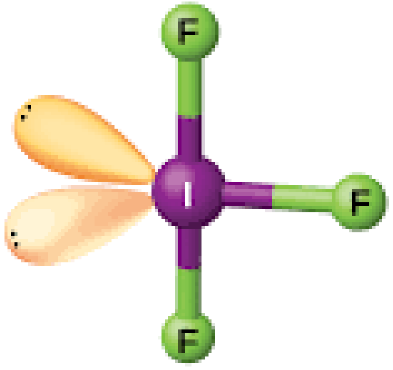
Hybridisation structure =



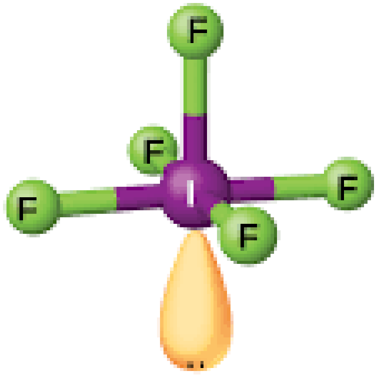
Hybridisation shape = Square pyramidal

STRUCTURE OF IF5

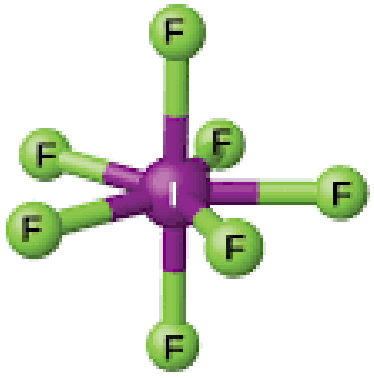




IF_3



IF_5



IF_7

Thank you

Interhalaogen compounds

INTERHALOGEN COMPOUNDS

The binary covalent compounds formed by halogens amongest themselves are known as interhalogen compounds. They have the general formula, AX_n where A is a large halogen atom and X is a smaller halogen atom. The value of n is 1,3,5 and 7. Some examples are given in the table.

S.No.	Type	Example
1.	AX	ClF, BrF, BrCl, ICl, IBr
2.	AX_3	ClF_3, BrF_3, IF_3, ICl_3
3.	AX_5	ClF_5, BrF_5, IF_5
4.	AX_7	IF_7

(i) In naming the compound, the less electronegative element is mentioned first.

4.	AX_7	IF_7

(i) In naming the compound, the less electronegative element is mentioned first.

(ii) All interhalogen compounds are diamagnetic, since **n** is odd.

(iii) Iodine alone forms all types of interhalogen compounds mainly due to its big size and electropositive nature.

(iv) It is found that the greater the difference in the electronegativity values of the two halogens, the stronger is the bond between them.

We shall discuss few interhalogen compounds in detail.

Interhalogens of AX type

Iodine monochloride

Preparation. (*i*) It can be prepared by the direct combination of liquid chlorine with solid iodine under controlled conditions of temperature and pressure.

$$\underset{\text{Solid}}{I_2} + \underset{\text{Liquid}}{Cl_2} \longrightarrow 2ICl$$

(*ii*) It may also be obtained by heating iodine with potassium chlorate.

$$KClO_3 + I_2 \longrightarrow KIO_3 + ICl$$

Properties. (*i*) It is a reddish oil and is readily hydrolyzed by water.

$$ICl + H_2O \rightleftharpoons HCl + HOI$$

(*ii*) It is also decomposed by sodium hydroxide.

$$ICl + NaOH \longrightarrow NaCl + HOI$$

(*iii*) When electricity is passed through iodine monochloride, iodine is liberated at the cathode and chlorine at the anode, indicating that the compound is a chloride of monovalent positively charged iodine.

$$2ICl \longrightarrow I^+ + [ICl_2]^-$$

(*iv*) With alkali metal halides it gives polyhalides.

$$KCl + ICl \longrightarrow KICl_2$$

$KCl + ICl \longrightarrow KICl_2$

Uses. (*i*) It is used to iodinate organic compounds, though chlorination may also occur depending on the condition of reaction.

$$\text{Salicylic acid} \xrightarrow{\text{ICl vapour}} \text{Chlorination}$$

$$\text{Salicylic acid} \xrightarrow[\text{nitrobenzene}]{\text{ICl in}} \text{Iodination}$$

(*ii*) It is used as Wij's reagent in the estimation of iodine number which is a measure of unsaturatio[n] of fats and oils.

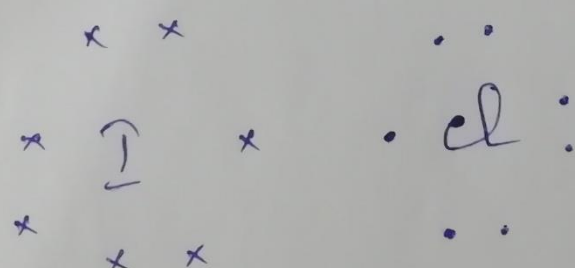
$$-CH=CH- + ICl \longrightarrow -\underset{\displaystyle I}{\underset{|}{CH}}-\underset{\displaystyle Cl}{\underset{|}{CH}}-$$

Iodine Monochloride ICl

Structure

I_{53} $[Kr]_{36}\ 4d^{10}\ 5s^{2}\ 5p^{5}$

Cl_{17} $1s^{2}\ 2s^{2}\ 2p^{6}\ 3s^{2}\ 3p^{5}$



Linear structure.

he.

er.

Bromine trifluoride, BrF_3

Preparation. It is obtained by mixing bromine and fluorine vapours or by the action of ClF_3 on bromine at 238 K.

$$Br_2 + 3F_2 \longrightarrow 2BrF_3$$

$$2ClF_3 + Br_2 \longrightarrow 2BrF_3 + Cl_2$$

Properties. (*i*) Bromine trifluoride is a greenish-yellow fuming liquid having b.p. 400 K.

(*ii*) With alkali fluorides it forms polyhalides.

$$KF + BrF_3 \longrightarrow KBrF_4$$

(*iii*) It converts many metal oxides, halides and even metals to their respective fluorides.

$$6MO + 4BrF_3 \longrightarrow 6MF_2 + 2Br_2 + 3O_2$$

where M is a divalent metal.

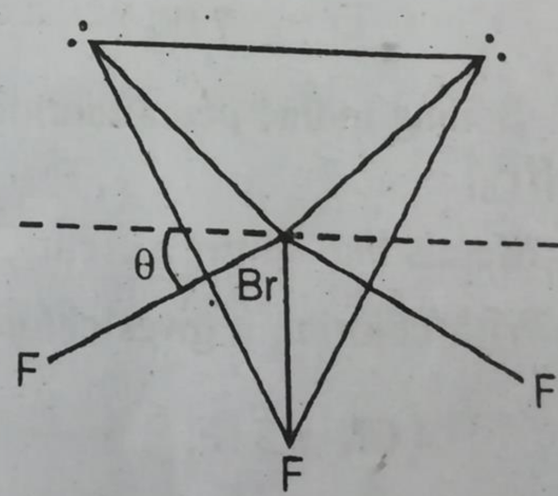
(*iv*) It ionizes considerably as

$$2BrF_3 \rightleftharpoons BrF_2^+ + BrF_4^-$$

... involves sp^3d hybridisation. The T shape structure can b

Structure

In bromine trifluoride molecule, ten electrons are present in the valency shell of the bromine atom. Out of these, two pairs are lone pairs. The expected geometry is trigonal bipyramidal. But due to the presence of the two lone pairs, the shape is distorted. To avoid repulsion, the bonding pairs contract and it has a *bent T shape*. It makes use of sp^3d - hybridisation.



IODINE PENTAFLUORIDE, IF_5

Preparation

(a) It is prepared by passing fluorine (diluted with nitrogen) over iodine in a heated quartz tube :

$$I_2 + 5\,F_2 \longrightarrow 2\,IF_5$$

(b) It is formed by the action of fluorine on I_2O_5 :

$$2\,I_2O_5 + 10\,F_2 \longrightarrow 4\,IF_5 + 5\,O_2$$

(c) It can also be obtained by heating I_2 with silver fluoride :

$$I_2 + 10\,AgF \longrightarrow 2\,IF_5 + 10\,Ag$$

Properties. (*i*) Liquid FI_5 conducts electricity and self-ionizes.

$$2IF_5 \rightleftharpoons IF_4^+ + IF_6^-$$

(*ii*) It undergoes hydrolysis as

$$IF_5 + 3H_2O \longrightarrow 5HF + HIO_3$$

(*iii*) Complex salts are formed by heating IF_5 with alkali halides.

$$KI + IF_5 \longrightarrow K[IF_6]$$

Structure of iodine pentafluoride

IF_5 is formed by sp^3d^2 hybridisation and hence is octahedral in structure as shown below (Fi

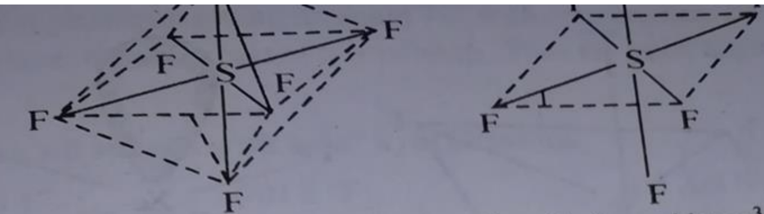


Fig. 3.37. Octahedral molecule of sulphur hexafluoride formed by sp^3d^2 hybridisation.

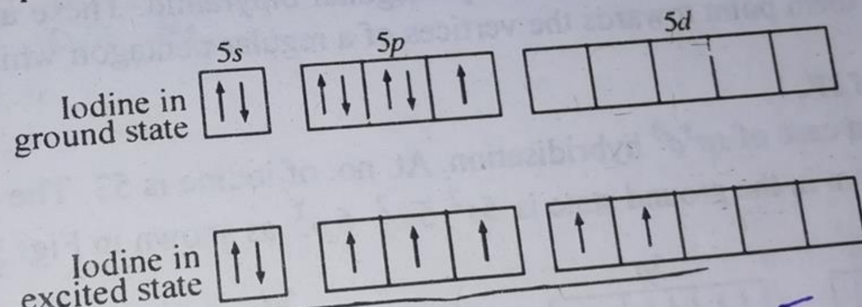
(ii) Shape of iodine pentafluoride IF_5.

It is a case of sp^3d^2 hybridisation. At. no. of iodine is 53. Its outer electronic configuration is

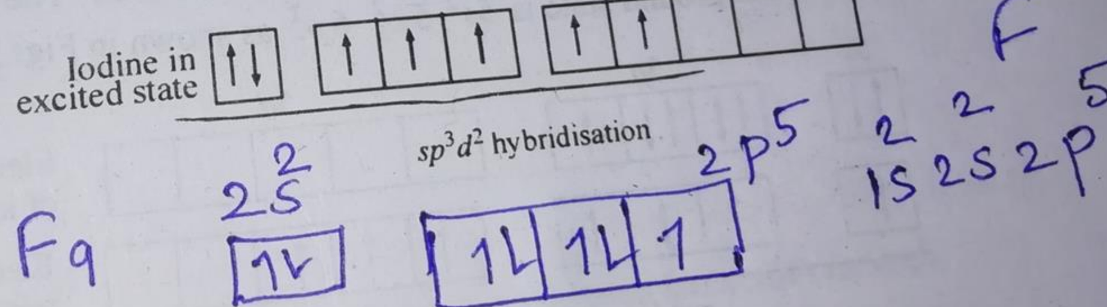
$$5s^2\ 5p_x^2\ 5p_y^2\ 5p_z^1$$

In the excited state two electrons from $5p$ orbitals are excited to $5d$ orbitals, thereby giving five single electrons for the formation of five bonds with fluorine. $5s$, $5p_x$, $5p_y$, $5p_z$ and two d-orbitals (sp^3d^2) hybridise to give equivalent orbitals oriented octahedrally.

However one position is occupied by a lone pair. Therefore actual shape is square pyramidal (Fig. 3.38).



sp^3d^2 hybridisation



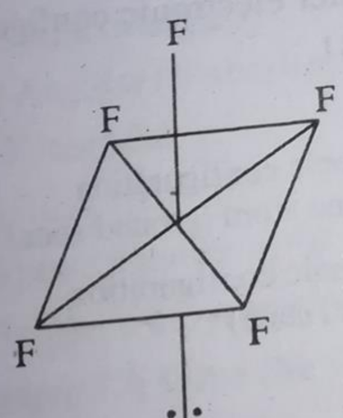


Fig. 3.38. Shape of IF_5 molecule.

...ic configuration in the

b) It is also formed by heating iodine with potassium chlorate.

$$KClO_3 + I_2 \longrightarrow KIO_3 + ICl$$

Properties

It is a dark red liquid but solidifies on standing at 290 - 300 K. It exists in two forms -- stable and metastable form.

It is readily hydrolysed forming hypoiodite and iodate ions.

$$Cl^- + OI^- \xleftarrow{H_2O} ICl \xrightarrow{H_2O} Cl^- + IO_3^-$$

When current is passed through ICl, iodine is liberated at the cathode and chlorine at the anode.

Alkali metal chlorides like KCl dissolve in it.

$$KCl + ICl \longrightarrow K^+ [ICl_2]^-$$

It is used to iodinate organic compounds. Chlorination may also occur depending on the condition.

$$\text{Chlorination} \xleftarrow[\text{vapour}]{ICl} \text{Salicylic acid} \xrightarrow[\text{nitrobenzene}]{ICl \text{ in}} \text{Iodination}$$

Uses

(1) Liquid I Cl is used as an ionising solvent.

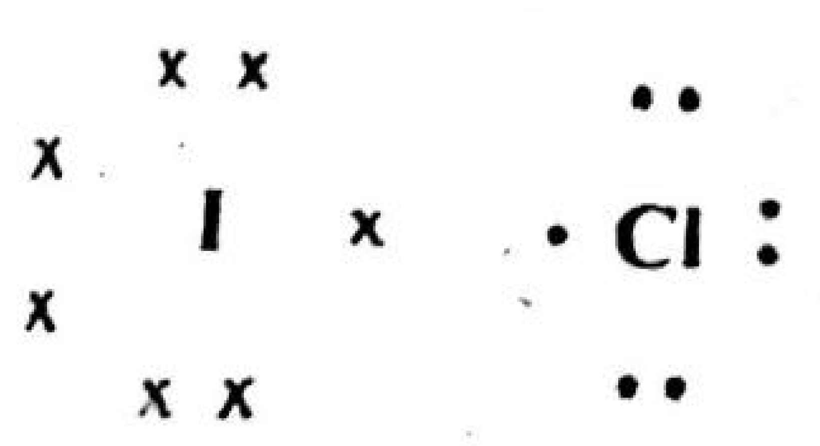
(2) It finds use as *Wij's reagent* in the estimation of iodine value of fats and oils.

$$-HC=CH- + ICl \longrightarrow -\underset{\displaystyle I}{\underset{|}{CH}}-\underset{\displaystyle Cl}{\underset{|}{CH}}-$$

(3) It is used to iodinate organic compounds.

Structure

It has the following linear structure :



BROMINE TRIFLUORIDE, BrF_3

Preparation

(a) It is prepared by mixing bromine vapour and fluorine in a stream of nitrogen :

$$Br_2 + 3F_2 \longrightarrow 2BrF_3$$

(b) It is also formed by the action of ClF_3 on bromine at 283 K :

$$ClF_3 + Br_2 \longrightarrow BrF_3 + BrCl$$

chlorine tri fluoride

Properties

It is a colourless liquid (b.p. 400 K). It has got high electrical conductivity.

It undergoes auto-ionisation as –

$$2BrF_3 \rightleftharpoons BrF_2^+ + BrF_4^-$$

It reacts with metals and metallic oxides forming salts containing BrF_2^+ and BrF_4^- groups. For example,

$$[BrF_2]^+ [SbF_6]^- ; [BrF_2]^+ [AuF_4]^-$$

$$[BrF_2]_2^{+2} [SnF_6]^{2-} \text{ and } K^+ [BrF_4]^-$$

It converts many metal oxides and chlorides into fluorides as it is a good fluorinating agent :

$$6M + 2BrF_3 \longrightarrow 6MF + Br_2$$

$$3MCl + BrF_3 \longrightarrow 3MF + BrCl_3$$

It is decomposed violently by water giving bromine, oxygen, bromic acid ($HBrO_3$) and hydrofluoric acid :

$$3BrF_3 + 5H_2O \longrightarrow O_2 + Br_2 + HBrO_3 + 9HF$$

It displaces oxygen from oxides like SiO_2, CuO etc.

$$3SiO_2 + 4BrF_3 \longrightarrow 3SiF_4 + 2Br_2 + 3O_2$$

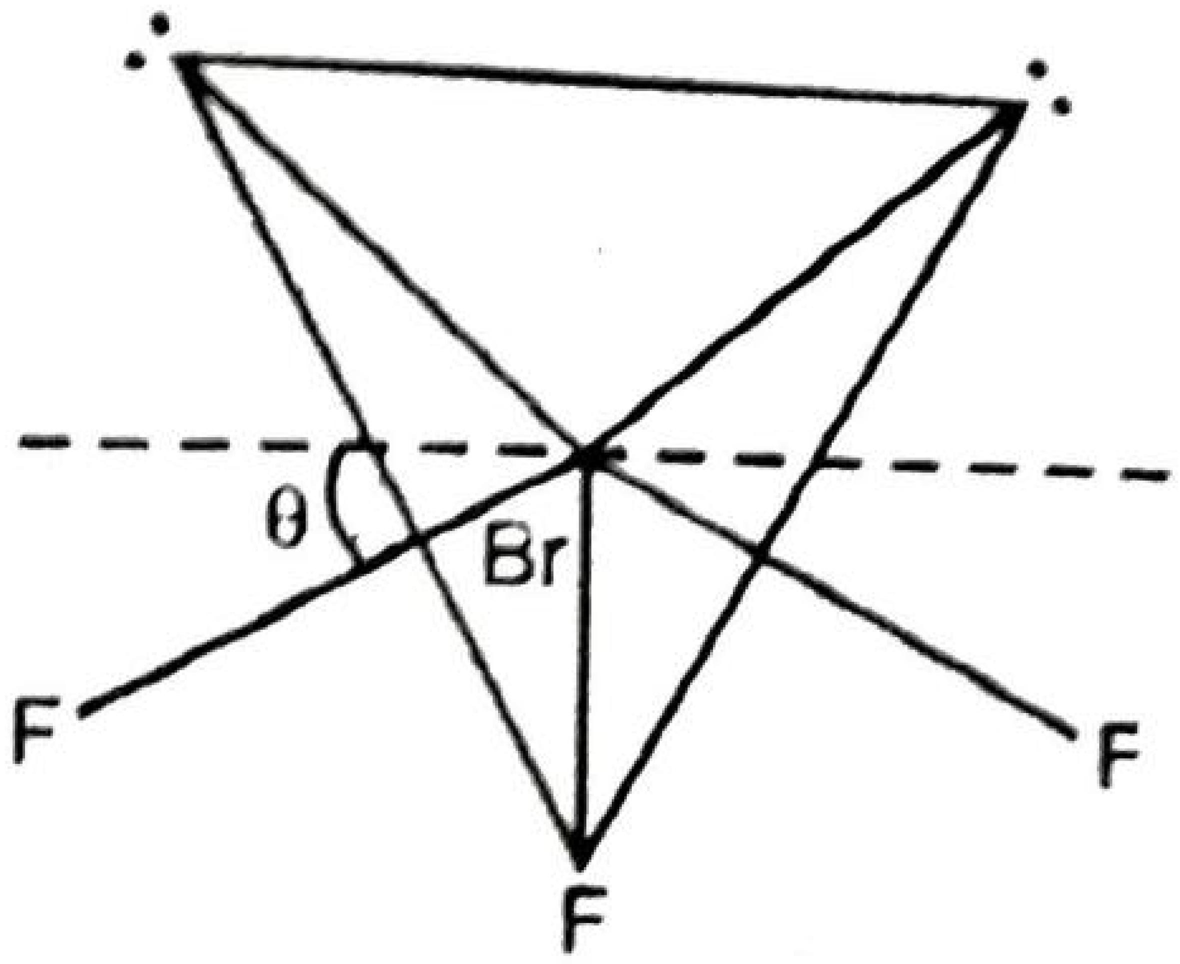
Uses

(1) BrF_3 is used as a solvent.

(2) It is a good fluorinating agent.

Structure

In bromine trifluoride molecule, ten electrons are present in the valency shell of the bromine atom. Out of these, two pairs are lone pairs. The expected geometry is trigonal bipyramidal. But due to the presence of the two lone pairs, the shape is distorted. To avoid repulsion, the bonding pairs contract and it has a *bent T shape*. It makes use of sp^3d - hybridisation.



IODINE PENTAFLUORIDE, IF_5

Preparation

(a) It is prepared by passing fluorine (diluted with nitrogen) over iodine in a heated quartz tube :

$$I_2 + 5\,F_2 \longrightarrow 2\,IF_5$$

(b) It is formed by the action of fluorine on I_2O_5 :

$$2\,I_2O_5 + 10\,F_2 \longrightarrow 4\,IF_5 + 5\,O_2$$

(c) It can also be obtained by heating I_2 with silver fluoride :

$$I_2 + 10\,AgF \longrightarrow 2\,IF_5 + 10\,Ag$$

Properties

It is a colourless liquid (b.p. 330 K). It is a good conductor of electricity as it undergoes self ionisation as

$$2\,IF_5 \longrightarrow IF_4^+ + IF_6^-$$

It reacts with water to form iodic and hydrofluoric acids :

$$2\,IF_5 + 5\,H_2O \longrightarrow 10\,HF + I_2O_5$$

$$I_2O_5 + H_2O \longrightarrow 2\,HIO_3$$

When heated to about 773 K it decomposes to iodine and iodine heptafluoride.

$$7\,IF_5 \longrightarrow 5\,IF_7 + I_2$$

Boiling iodine pentafluoride dissolves potassium fluoride to give, $K^+[IF_6]^-$

Metals like silver, mercury, iron, copper are slowly attacked by IF_5.

With chlorine it gives chlorine trifluoride.

$$4\,Cl_2 + 3\,IF_5 \longrightarrow 5\,ClF_3 + 3\,ICl$$

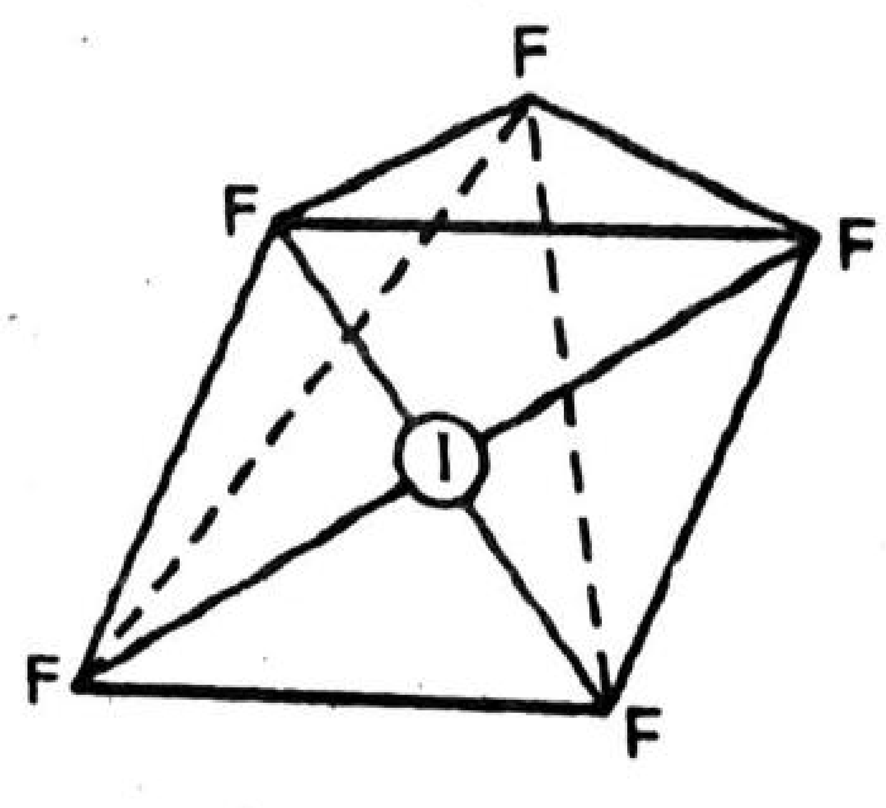
The alkali metals, sulphur, bromine react when heated.

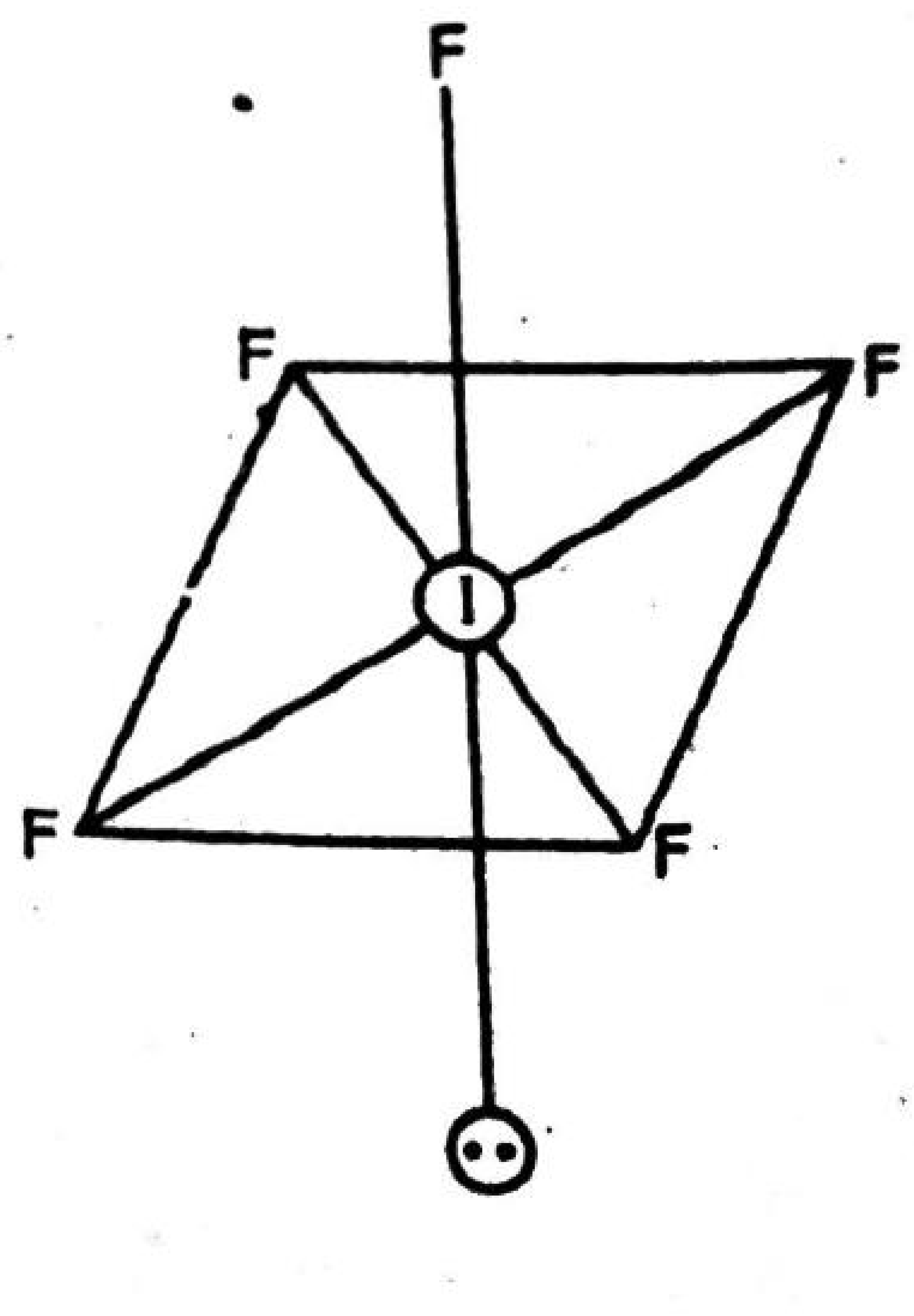
Uses

IF_5 is used as an ionising solvent.

Structure

There are six pairs of electrons around the central atom, iodine. Of these one is a lone pair. Therefore, the expected geometry of octahedron is slightly distorted and the molecule has a square pyramidal structure.





Chemistry of iodine heptafluoride, IF_7

Preparation

(i) It can be prepared by passing flourine over IF_3 or IF_5 at 300°C.

$$IF_3 + 2F_2 \xrightarrow{300°C} IF_7$$

$$IF_5 + F_2 \xrightarrow{300°C} IF_7$$

(ii) It can be prepared from the following reaction.

$$PbI_2 + 8\,F_2 \longrightarrow PbF_2 + 2\,IF_7$$

(iii) It can also be prepared by passing excess fluorine over gaseous iodine at 300° C

$$I_2 + 7\,F_2 \longrightarrow 2IF_7$$

Properties

It is a colourless gas with m.p.6.5°C. It is diamagnetic in nature.

IF_7 is a stronger fluorinating agent than IF_5. It reacts with most elements in cold or on warming including glass or silica at 100° C

$$Si\,O_2 + 2\,IF_7 \rightarrow Si\,F_4 + 2\,IOF_5$$

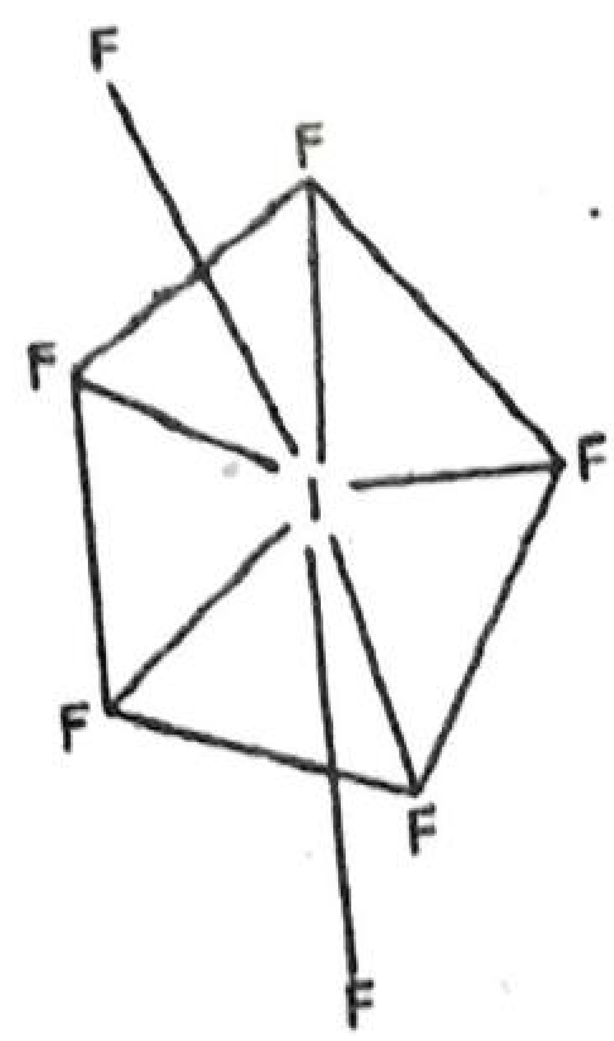
Vapours of IF_7 undergo smooth hydrolysis to HIO_4 and HF.

$$IF_7 + 4\,H_2\,O \rightarrow HI\,O_4 + 7\,HF$$

An oxofluoride (IOF_5) is formed with less amount of water at room temperature.

$$IF_7 + H_2\,O \rightarrow 2\,HF + IOF_5$$

It reacts with most metals (except the platinum group), non metals including Cl_2 and I_2, water (to give periodate and fluoride ions), many inorganic compounds and a variety of organic materials.



It explodes when heated with hydrogen.

Two of seven fluorine atoms present in IF_7 can easily be removed.

Uses

It is used as a fluorinating agent.