Chemical Bonding

- Covalent bond
- ionic bond
- metallic bond
- dative bond or co ordinate bond
- Hydrogen Bond

Covalent Bonding

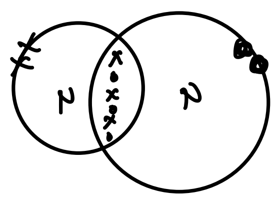
- Covalent bond is formed due to the share of electrons between two non-metals.
- shared electron pair is attracted by two nuclei.
- The attraction is electrostatic

Molecule	Displayed Formula	No. of bond pair	No. of lone pair	Total number of electrons in the molecule
H_2	H−H	1	0	2
O_2	O=O	2	4	16
N_2	N≡N	3	2	14
N_2H_4	H_2N-NH_2	5	2	18
CO_2	O=C=O	4	4	22
H_2O_2	H−O−O−H	3	4	18

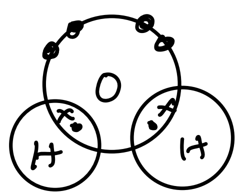
[illegible]				
CH_3OH	$H-\overset{H}{\underset{H}{C}}-O-H$	5	2	18
CH_3^+	$H-\overset{+}{\underset{H}{C}}-H$	3	0	8
CS_2	$S=C=S$	4	4	38
OH^-	^-O-H	1	3	10

Dot and Cross diagram

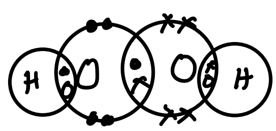
① N_2



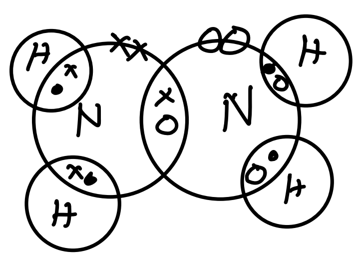
② H_2O



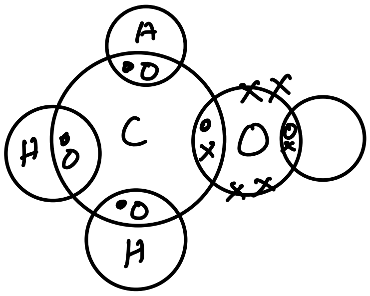
③ H_2O_2

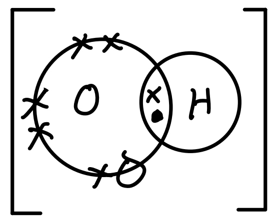


④ N_2H_4

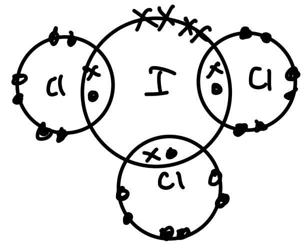


⑤ CH_3OH

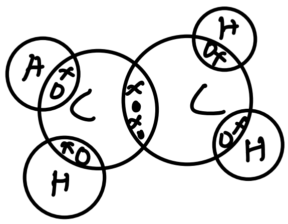




(8)



(6) C_2H_4



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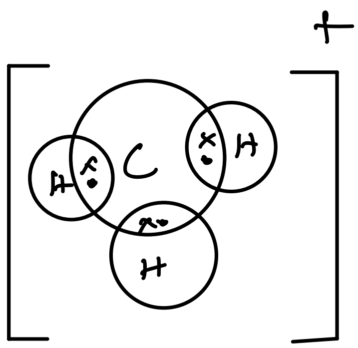
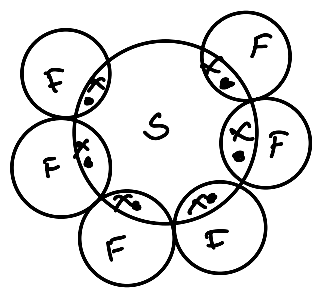




SF_6

IF_7 is possible but ICl_7 is not possible

- Chlorine has larger atomic radius than the atomic radius of fluorine

- Iodine atom can arrange seven fluorine atoms around itself.

- Iodine atom cannot arange seven

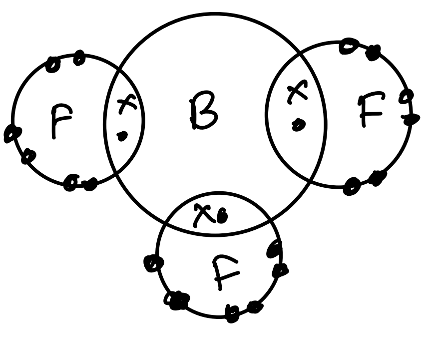
chlorine atoms around itself.

PCl_5 is possible but NCl_5 is not possible

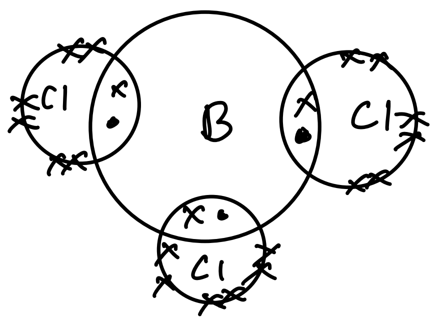
- second shell of nitrogen can contain maximum eight electrons.
- Third shell of phosphorous can contain maximum 18 electrons.

Electron deficient compounds

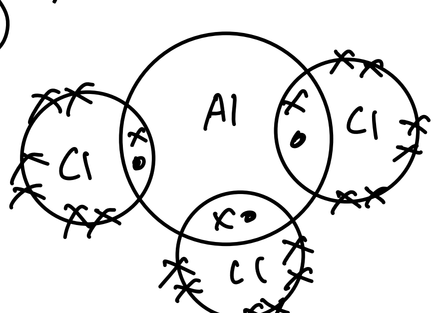
① BF_3



② BCl_3



③ $AlCl_3$

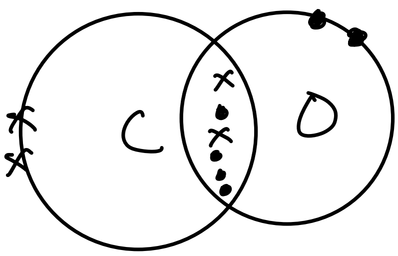


Dative bond / Coordinate bond

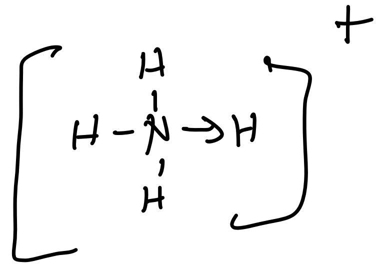
⊖ During the formation of a dative bond both electrons (shared pair electrons) are given from one atom and another atom share the electron pair.

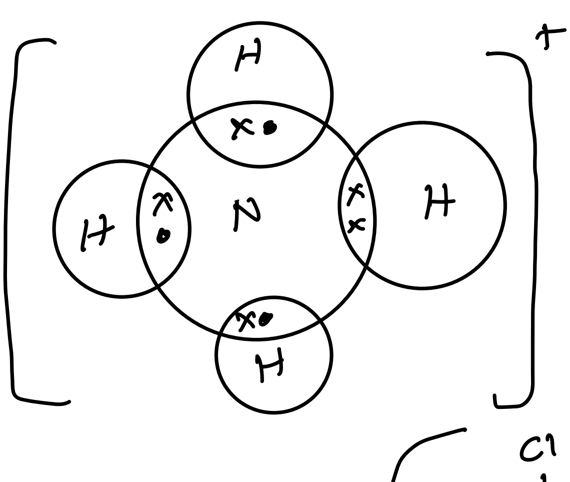
Examples

① CO

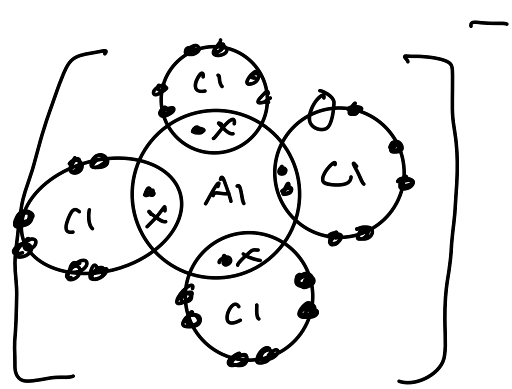


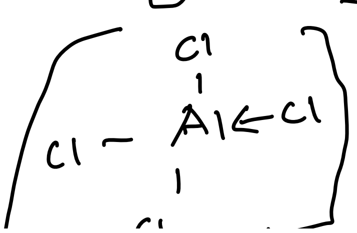
② NH_4^+



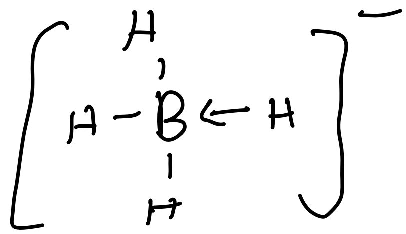


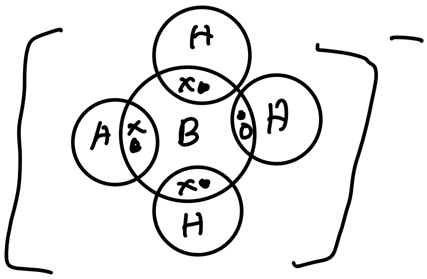
③ $AlCl_4^-$



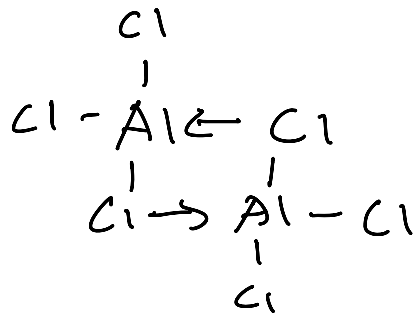


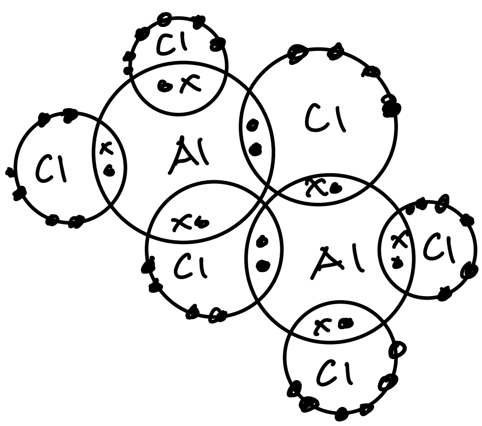
(4) BH_4^-



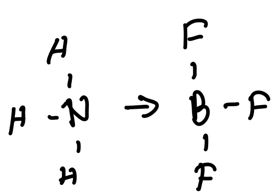


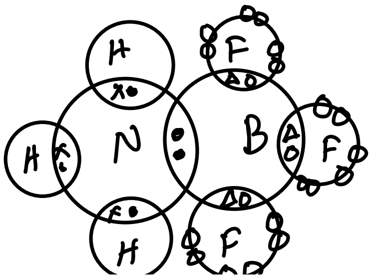
(5) Al_2Cl_6





$NH_3 + BF_3 \longrightarrow H_3N \rightarrow BF_3$

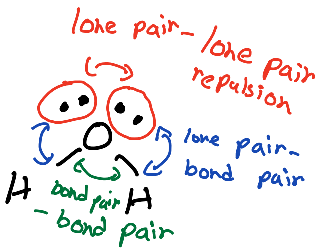




Shapes of the molecules by using VSEPR theory

VSEPR

Valence shell electron Pair Repulsion



Repulsion

- lone pair - lone pair repulsion
- lone pair - bond pair repulsion
- bond pair - bond pair repulsion

strength of repulsion decreases

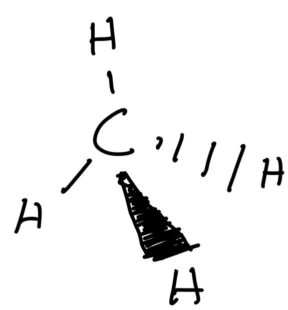
① Tetrahedral

Number of bond pair = four.

Number of lone pair = zero.

shape = Tetrahedral

Bond angle = 109.5^0



$SiCl_4$	SiH_4
$GeCl_4$	CCl_4
NH_4^+	BH_4^-

$AlCl_4^-$

② Linear

Number of bond pair = two.

Number of lone pair = zero

bond angle 180°

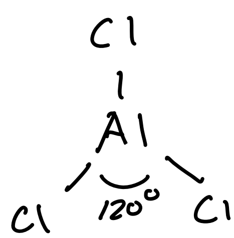
Cl – Be – Cl

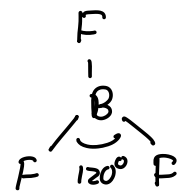
③ Trigonal plannar

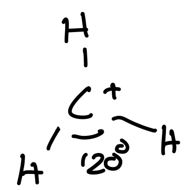
Three bond pair

No lone pair

Bond angle 120°







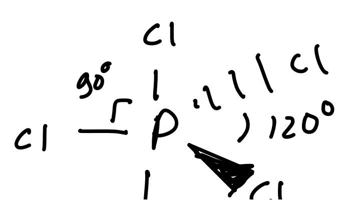
④ Trigonal bipyramidal

Number of bond pair five

Number of lone pair zero

Bond angle 90° and 120°.

PCl_5



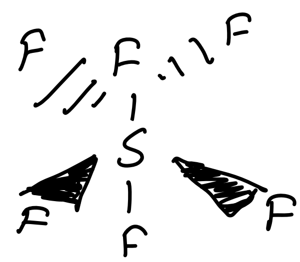
(5) Octahedral

Number of bond pair six

Number of lone pair zero

bond angle 90°

SF_6

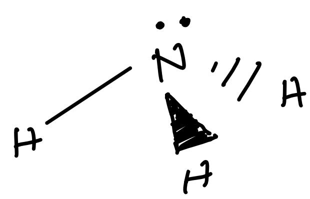


(6) Pyramidal

Number of bond pair = three

Number of lone pair = one

Bond angle = 107°



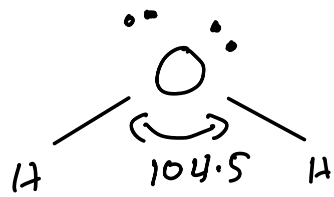
(7) V-shaped / bent / non-linear

Number of bond pair = two

Number of lone pair = two

shape = v-shaped/ bent

bond angle 104.5°

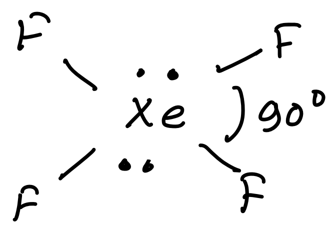


⑧ <u>Square Plannar</u>

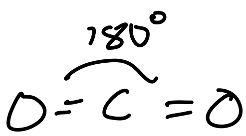
Bond pair = four

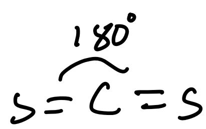
Lone pair = two

Bond angle = 90°



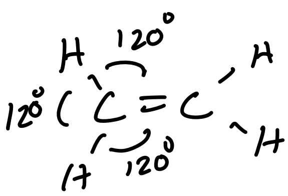
⑨

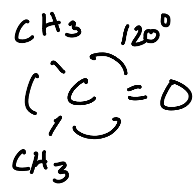




Linear

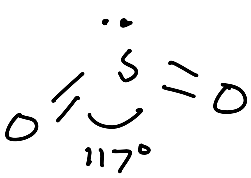
⑩



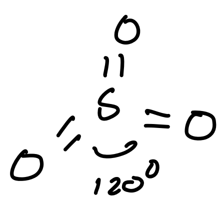


(plannar)

⑪



⑫



Trigonal plannar

⑬

$PH_3 = 90°$

Ionic bond

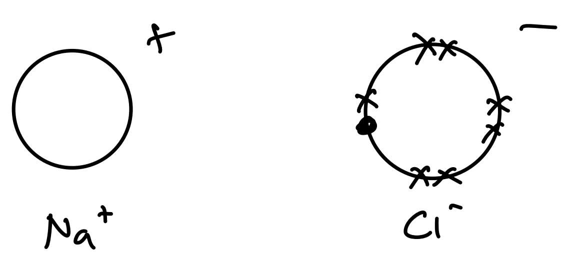
- Ionic bond is formed between metal and non metal
- Electrons are transferred from metal to non metal.
- Metallic atoms donate electrons to form cations.
- Non metallic atoms gain electrons to form anions.
- Ionic bond is formed due to the strong electrostatic attraction between the oppositely charged ions (cations and anions).

The strength of ionic bond depends on :

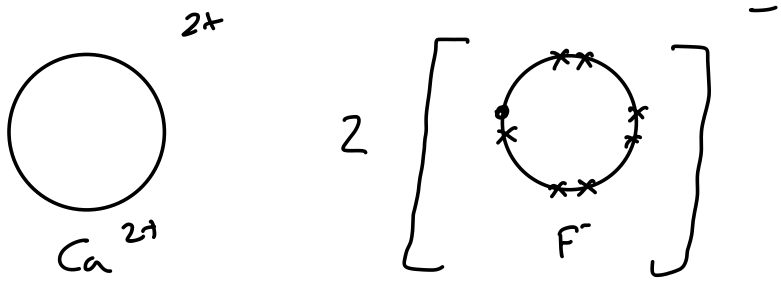
- charge on the ions.
- ionic radius
- charge density

- With the increase of charge on the ion, the strength of electrostatic attraction increase.
- with the increase of charge density the strength of ionic bond increase.

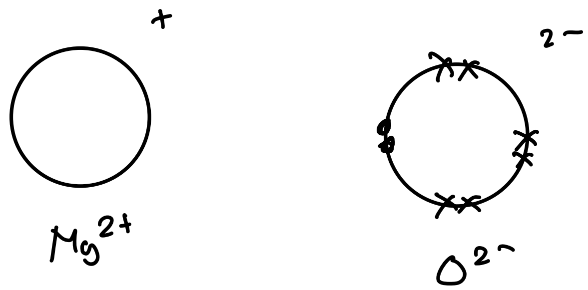
Examples : NaCl



CaF_2

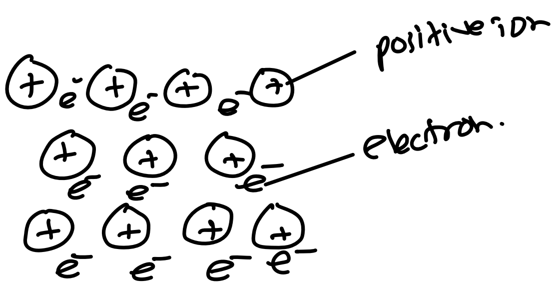


MgO



Metallic Bond

The bond which is formed due the strong electrostatic force attraction between positively charged metal ions and (sea of) delocalised electrons is called metallic bond.



Strength of ionic bond depends on:
- Charge on the ions
- ionic radius
- Number of delocalised electrons

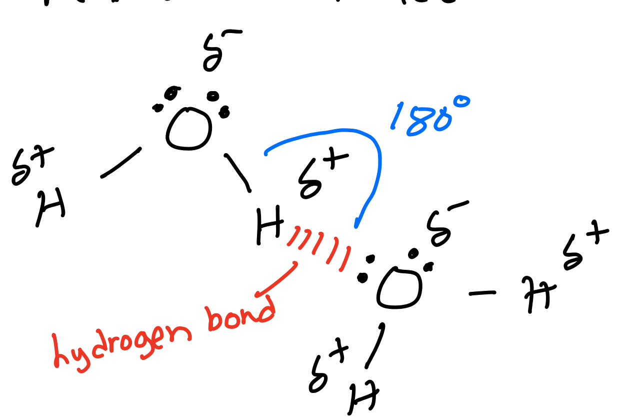
⊖ With the increase of charge on the ions/ Number of delocalised electrons, the strength of metallic bond increase.

⊖ With increase of ionic radius of the positive metal ions, the strength of metallic bond decrease.

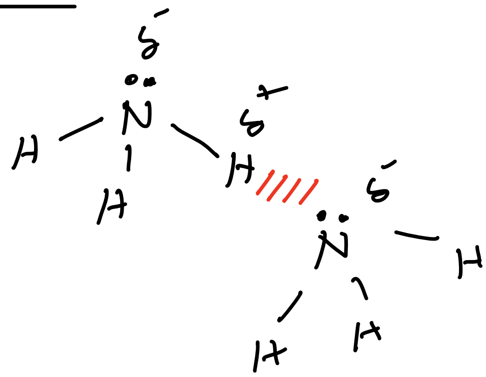
Hydrogen bond

⊖ The bond which is formed between hydrogen atom in one molecule and oxygen/ nitrogen/ Fluorine atom of another molecule is called hydrogen bond.

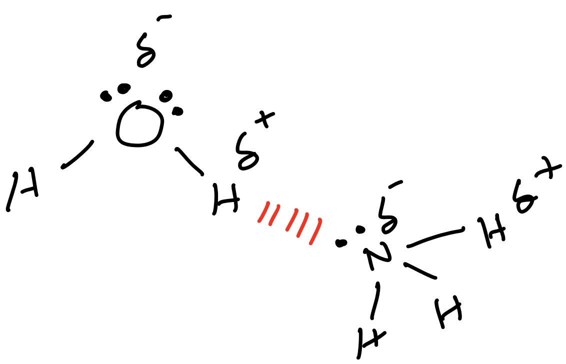
example: water molecule



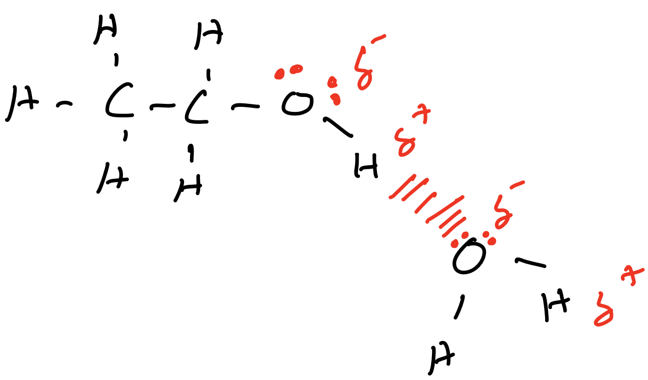
Ammonia



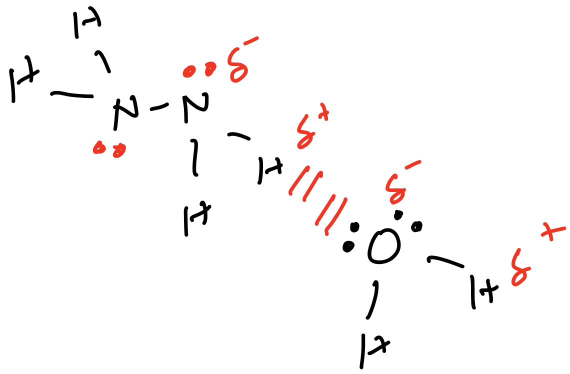
Water and Ammonia



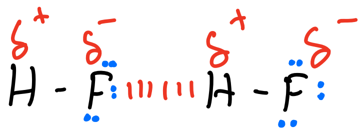
Ethanol and Water



Hydrazine and water



Two Hydrogen Fluoride molecules



Structure and Physical Properties

- Structure
  - Simple Covalent or Simple Molecular
  - Giant Covalent or Giant Molecules
  - Giant ionic lattice
  - Giant metallic
  - Simple atomic

Giant ionic lattice Physical Properties

① Ionic compounds have high m.p and b.p.

Explanation

- Giant ionic lattice
- strong ionic bonds
- High heat energy is needed to break the strong electrostatic attraction between the oppositely charged ions.

Example-1

NaCl has high mp and bp.

- NaCl has giant ionic lattice and strong ionic bonds.
- High heat energy is needed to break the strong ionic bonds(strong electrostatic force of attraction between the oppositely charged ions.

Example-2

MgO and Na_2O

- MgO has higher m.p than Na_2O
- Magnesium ion has smaller ionic radius and greater charge than sodium ion.
- Magnesium ion has greater charge density than sodium ion.

- Stronger electrostatic attraction between magnesium ion and oxide ion.

NaCl has higher m.p than RbCl

- Sodium ion has smaller ionic radius than Rb^+ ion.
- Na^+ ion has greater charge density than Rb^+ ion.
- Stronger electrostatic attraction between Na^+ and Cl^- ions.

② Conductivity

- Ionic compounds are non conductor in the solid state
- Ionic compounds are conductors in the liquid state (molten state)
- Soluble ionic compounds conduct electricity in aqueous.

* Explanation

- In the solid state ionic compounds have no free ions to move to conduct electricity.

⊖ In the liquid state (molten state) ionic compounds have free ions to move to conduct electricity.

Example - 1

⊖ NaCl can not conduct electricity in the solid state at (r t p).

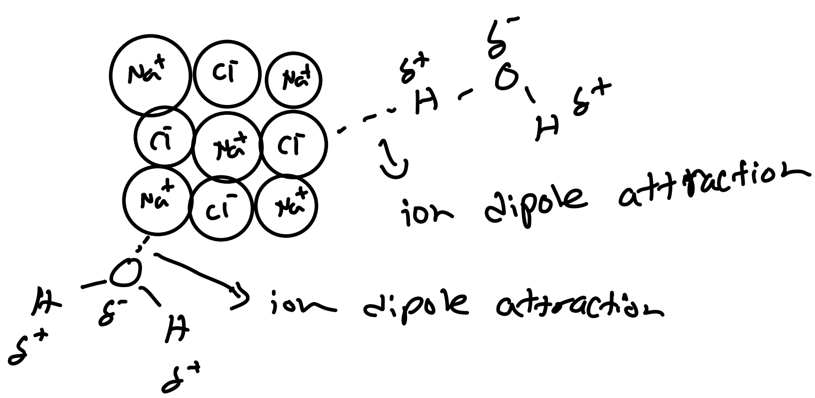
⊖ NaCl can conduct electricity in aqueous

⊖ NaCl can conduct electricity in the molten state.

③ Solubility

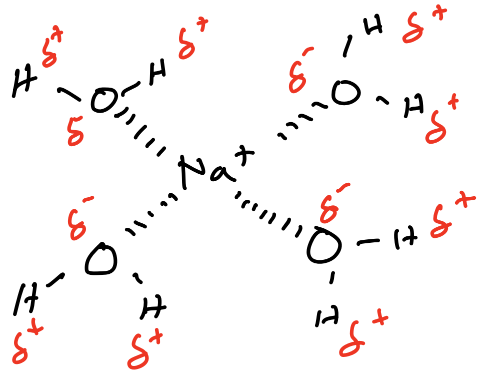
⊖ Most of the ionic compounds are soluble in polar solvent (H_2O).

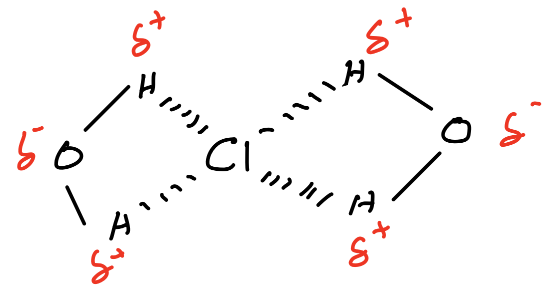
⊖ Ionic compounds are insoluble in non-polar solvent.



⊖ Sodium ions are attracted by the partial negative charge oxygen in water.

⊖ Chloride ions are attracted by the partial positive charge hydrogen in water.





④ <u>Physical state</u>

Ionic compounds are existed as a solid at rtp.

<u>Explanation</u>

⊖ Ionic compounds have oppositely charged ions.

⊖ Due to the strong electrostatic force of attraction, ions are closely packed and are existed as a solid at rtp.

<u>Physical properties of the Giant metallic lattice</u>

① Most of the metals have high m.p and b.p (Gr-1 metals, mercury have low m.p)

<u>Explanation</u>

⊖ Giant metallic

⊖ Strong electrostatic attraction between positively charged metal ions and delocalised electrons.

- High heat energy is needed to break/overcome the strong electrostatic attraction.

Example-1

Na Mg Al

$\longrightarrow$

m.p increase - Explain

- Charge on the ions increase.
- Number of delocalised electrons increase.
- Ionic radius decrease.
- The strength of electrostatic attraction between positively charged metal ions and delocalised electrons increase.
- Metallic bond becomes stronger.
- For Al highest heat energy is needed to break the electrostatic attraction.

Li
Na
K
Rb
Cs
$\downarrow$ mp decrease

- Charge on the ions identical
- Number of delocalised electrons identical
- ionic radius increase

(+) The strength of electrostatic force of attraction between the positively charged metal ions and delocalised electrons decrease.

(−) Metallic bond becomes weaker

(+) Less energy is needed to break the metallic bond.

(2) Metals can conduct electricity in the solid state at r.t.p.

Metals can conduct electricity in the molten state (liquid state)

Explanation

(−) Metals have delocalised electrons in the solid state / liquid state.

(−) Electrons can move to conduct electricity

(3) Metals are ductile and malleable.

(−) Metals have layers of positively charged metal ions.

(−) Layers can slide over each other when force is applied.

Physical properties of the simple molecular/ simple covalent substances

Types of intermolecular forces of attractions

① Van der Waal's forces of attraction

or

Induced dipole – Induced dipole attraction

or

Temporary dipole – Induced dipole

or

Instantaneous dipole – Induced dipole

or

dispersion force

② Permanent dipole – dipole attraction.

③ Hydrogen bond

$O=C=O$

(Non – polar)

$\overset{\delta^+}{C}\equiv\overset{\delta^-}{O}$

polar

① Simple molecular substances have low m.p and b.p

Explanation

To melt or boil simple molecular substances we have to break the attraction between two molecules.

Example : F_2 Cl_2 Br_2 I_2

$$\xrightarrow[\text{m.p increase}]{}$$

⊖ From F_2 to I_2 number of electrons increase.

⊖ The strength of van der Waal's force of attraction increase.

⊖ More energy is needed to break the attraction between two molecules.

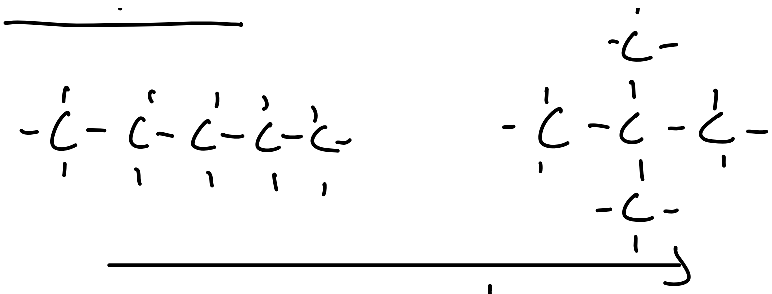
Example: Phosphorous , Sulfur , Chlorine
P_4 S_8 Cl_2

⊖ molecular formula
P_4 S_8 Cl_2

⊖ From Cl_2 to S_8 number of electrons increase.

⊖ The strength of van der Waal's forces of attraction increase.

Example-3



m.p decrease

- m.p decrease
- Contact area decrease
- The strength of van der Waal's forces of attraction decrease
- less energy is needed to overcome the attraction between two molecules.

Example-4

Carbon monoxide (CO)

Nitrogen (N_2)

- Carbon monoxide has higher b.p than Nitrogen

$$\overset{\delta+}{C} \equiv \overset{\delta-}{O} \text{ - - - - - - } \overset{\delta+}{C} \equiv \overset{\delta-}{O}$$

pd-pd attraction

$$\overset{\delta-}{N} \equiv \overset{\delta+}{N} \text{ ||||| } \overset{\delta-}{N} \equiv \overset{\delta+}{N}$$

id - id

⊝ Carbon monoxide is a polar molecule.
⊝ Nitrogen is a non-polar molecule.
⊝ Pd - Pd force of attraction is stronger than id - id force of attraction.

Example 5

Water has higher bp than carbon dioxide.
⊝ Carbon dioxide molecules have induced dipole - induced dipole attraction.
⊝ water molecules have permanent dipole, dipole attraction and hydrogen bonding.
⊝ Hydrogen bond is stronger than the induced dipole - induced dipole attraction.

② Conductivity

Non-conductor at r.t.p
Non conductor of electricity in the solid state
Non conductor of electricity in the liquid state.

** Simple molecular substances which can react with water to make free ions will conduct electricity in aqueous.

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

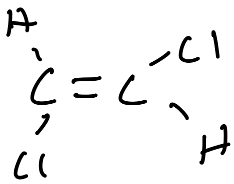
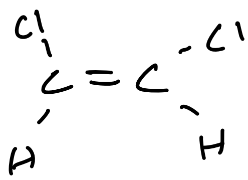
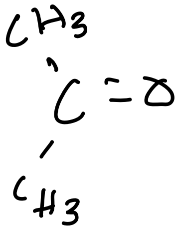
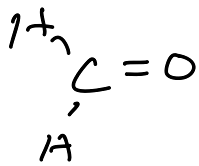
③ Solubility

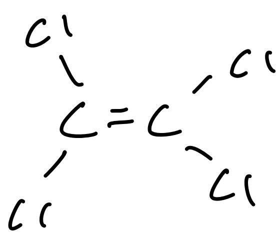
- Simple molecular substances are insoluble in polar solvent and soluble in non polar solvent.

- Alcohol, Amine, Carboxylic Acid, Ammonia are soluble in water due to the formation of hydrogen bond with water molecules.

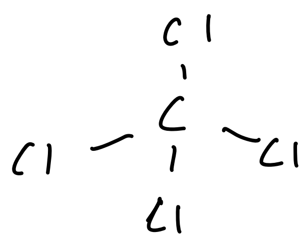
④ Physical State

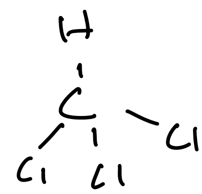
Simple molecular substances are solid, liquid, gas at r.t.p.

Molecule	Polar	Non-Polar
O=C=O		✓
C≡O	✓	
		✓
	✓	
	✓	
	✓	



Non Polar





Non polar Polar

Hybridization

Hybridization of Carbon

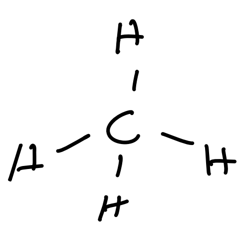
- sp^3
- sp^2
- sp

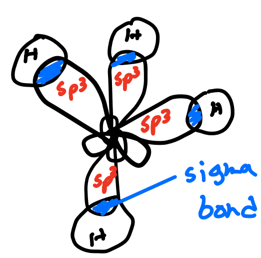
sp^3 hybridization

$C \rightarrow 1s^2 \quad 2s^2 \quad 2p_x^1 \quad 2p_y^1 \quad 2p_z$

$C^* \rightarrow 1s^2 \quad \underbrace{2s^1 \quad 2p_x^1 \quad 2p_y^1 \quad 2p_z^1}_{sp_3 \text{ hybridization}}$

(Four sp^3 hydrid orbitals)

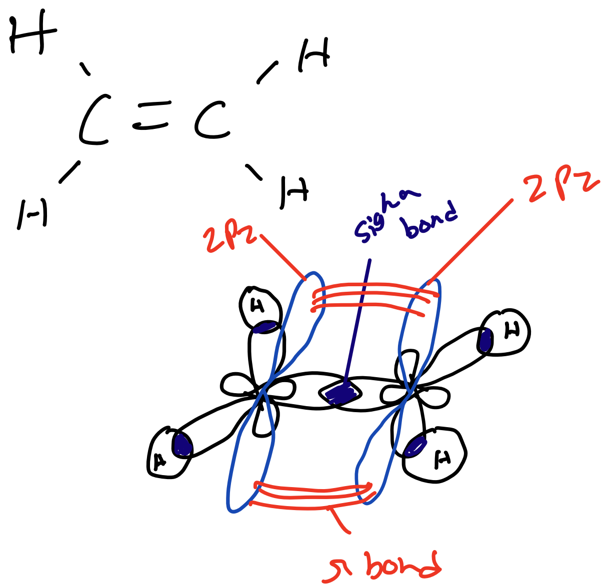




sp^2 hydridization

$$C^* \rightarrow 1s^2\ \underbrace{2s^1\ 2p_x^1\ 2p_y^1}_{sp^2 \text{ hydridization}}\ 2p_z^1$$

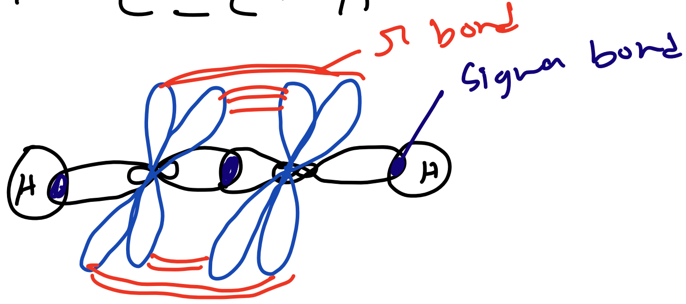
(Three sp^2 hybrid orbitals)



$\boxed{\text{sp hydridization}}$

$$C^* \rightarrow 1s^2\ \underbrace{2s^1\ 2p_x^1}_{sp \text{ hybridization}}\ 2p_y^1\ \ 2p_z^1$$

$H-C\equiv C-H$



Hybridization

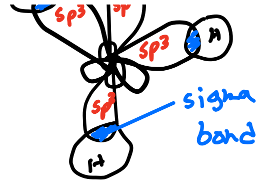
Hybridization is the mixing of atomic orbitals to form new orbitals with different energies and shapes than the original orbitals.

$\boxed{\text{sp}^3 \text{ hybridizations}}$

C^* (excited state) $\rightarrow 1s^2 \underbrace{2s^1\ 2p_x^1\ 2p_y^1\ 2p_z^1}_{4sp^3 \text{ hybrid orbitals}}$

The process by which one 2s orbital and three 2p orbitals are mixed with each other to make four new hydrid orbitals ($4sp^3$) is called sp^3 hydridization.

- sp^3 hybrid orbitals are pointing at $109.5°$ from one another.
- sp^3 hydridized carbon atom make four (σ) sigma bonds / single covalent bonds only.
- sp^3 hybridized carbon atoms can not make pi (π) bond



$\boxed{sp^2 \text{ hybridization}}$

$$\overset{*}{C} \rightarrow 1s^2\, \underbrace{2s^1\, 2p_x^1\, 2p_y^1}_{3\, sp^2 \text{ hybrid orbitals}}\, 2p_z^1$$

(excited state)

The process by which one 2s orbital and two 2p orbitals are mixed with each other to make three new hybrid orbitals is called sp^2 hybridization.

- sp^2 hybrid orbitals are pointing at 120° from one another.
- sp^2 hybridized carbon atoms can make three sigma bonds and one pi (π) bond.

<u>sp hydridization</u>

$$\overset{*}{C} \rightarrow 1s^2\, \underbrace{2s^1\, 2p_x^1}_{2sp \text{ hybrid orbitals}}\, 2p_y^1\, 2p_z^1$$

(excited state)

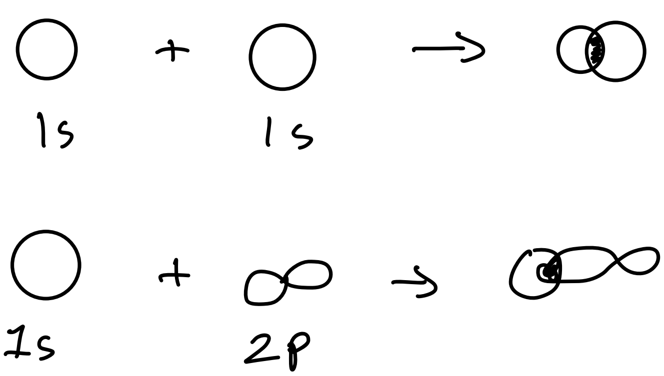
The process by which one 2s orbital and one 2p orbital are mixed together to make two new hybrid orbitals (2sp) is called sp hybridization.

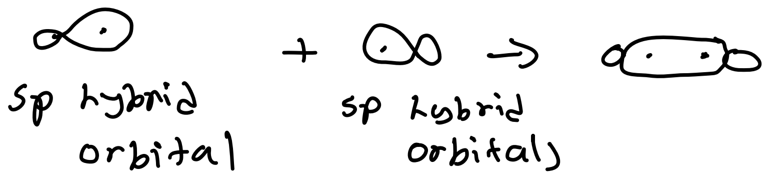
- sp hybrid orbitals are pointing 180° from one another.
- sp hybridized carbon atom can make two sigma bonds and two (π) bonds.

$\boxed{\text{Sigma bond}}$

The bond which is formed by head-on (end to end) overlapping between two orbitals is called sigma bond.

<u>Formation of sigma(σ) bond</u>

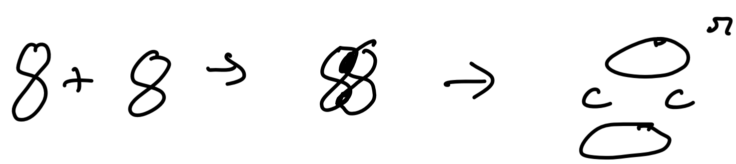




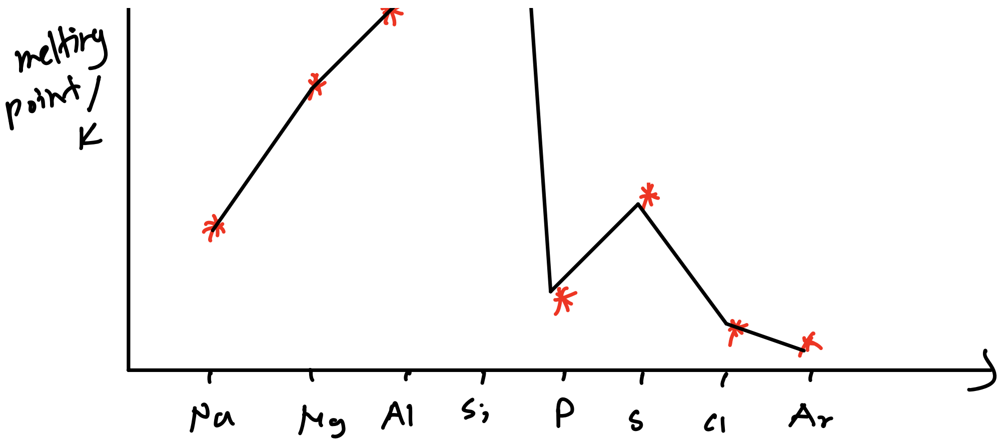
The formation of pi(π) bond

- The bond which is formed due to the sideways overlap of two p orbitals is called pi (π) bond.

- pi(π) bond has two areas of electron density. one above and one below the axis of joining the two atom centers.



The melting points of elements across the third period.



Increase in melting point from Na to Mg

- Increasing number of delocalised electrons
- Decreasing size of the cations.
- Increasing strength of attraction between cations (positively charged metal ions) and delocalised electrons.

The m.p increases more from Na to Mg than the increase from Mg to Al

⊖ Difference in size between Mg^{2+} and Al^{3+} is less than the difference in size between Na^{+} and Mg^{2+}.

⊖ Magnitude of increase in charge is less from +2 to +3 than +1 to +2.

The pattern of melting point from P to Ar

- Increase from P to S then decrease to Cl and Ar.
- General decrease from P to Ar with an increase from P to S.
- $S_8 > P_4 > Cl_2 > Ar$
- The greater the number of electrons in the molecule, the greater the strength of van der waal's force of attraction.
- melting point depends on strength of vander waal's force of attraction.

Silicon has the highest m.p

- Silicon has Giant covalent / Giant covalent structure.
- Many strong covalent bonds.
- More energy needed to break the strong covalent bonds.

Electronegativity

N O F

$H \rightarrow 2.1$

$C \rightarrow 2.5$

$F \rightarrow 4$ (3.98)

$Cl \rightarrow 3$ (3.16)

N → 3.0 (3.04)
O → 3.5 (3.44)
Na → 0.93

Br → 2.8 (2.96)
I → 2.5 (2.66)
At → 2.2
Si → 1.9 (1.8)

Predicting Bonding from Electronegativity

As an approximate indicator

- If the difference is less than 0.4, the bonding electrons are shared almost equally and the bond is mainly covalent.

- If the difference is between 0.4 and 2.0, the bond is polar covalent bond.

- If the difference is greater than 2.0 the bonding is mainly ionic.

Chloride	mp / °C	difference between the electronegativities
NaCl	801	2.2
$SiCl_4$	−69	1.3

Sodium Chloride ($NaCl$)

- Giant ionic lattice.
- Large difference in electronegativity.
- Transfer of electrons from sodium to chloride.
- Formation of ionic bonds.

Silicon tetrachloride ($SiCl_4$)

- Simple covalent / simple molecular.
- Smaller difference in electronegativity.
- sharing of electron pair to make covalent bonds.

polarity of C–H and Si–H bonds

$\overset{\delta^-}{C} - \overset{\delta^+}{H}$ $\qquad$ $\overset{\delta^+}{Si} - \overset{\delta^-}{H}$

CH_4 has no overall dipole moment

- Tetrahedral molecule
- Individual bond dipoles / partial charges cancel.

Bond energy and bond length

- Bond energy is the energy required to break one mole of a particular covalent bond in the gaseous state.
- Bond length measures the distance between the nuclei of two covalently bonded atoms.
- Bond length is the intermucsear distance of two covalently bonded ators.
- With the increase of atomic radius

  bond length increases
  bond strength decreases
  less energy is needed to break the bonds.

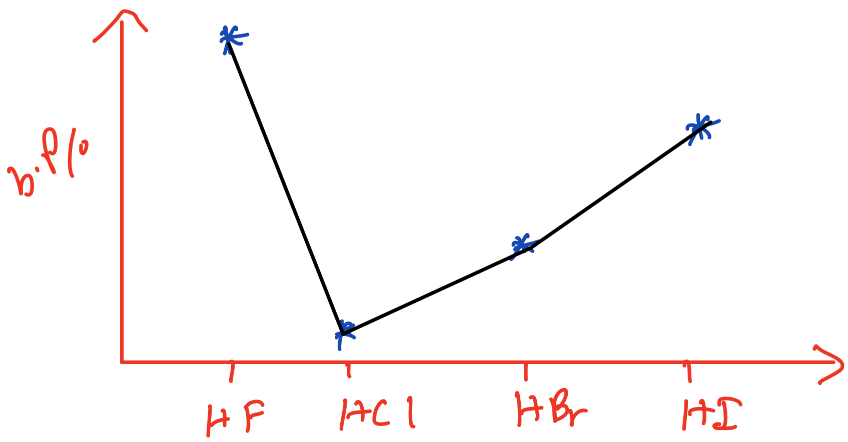
Thermal stability of the hydrogen halides

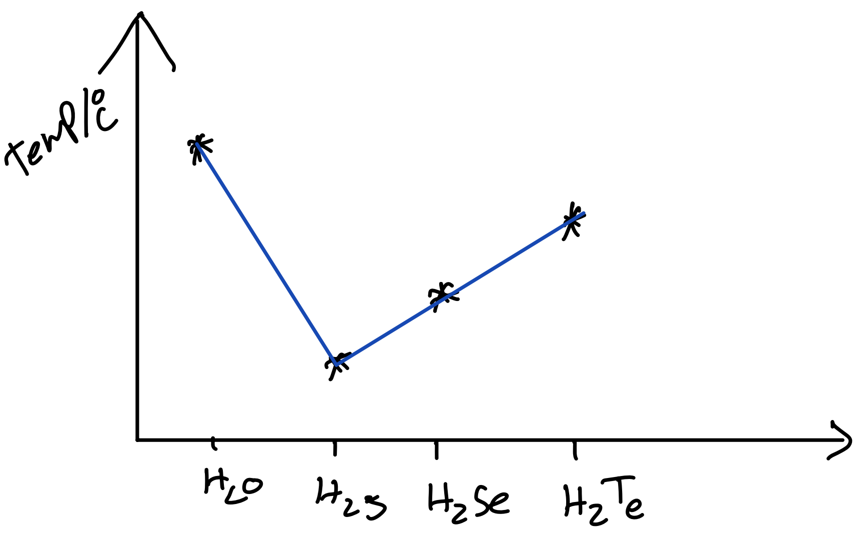
H - F
H - Cl
H - Br

- Thermal stability decreases
- Atomic radius of

$\ddot{H}-I$ ↓ the halogens increase.

- Bond length increase.
- Bond strength decrease.
- less energy is needed to break the bonds.





Surface Tension

- Surface tension is the tension caused by the attraction of the particles in the surface layer of a liquid.

- Water molecules on the surface can only be attracted sideways and downwards, but not upwards.

- This pulls the water molecule into the liquid and tightens the surface creating a surface tension.

Vaporisation

The change from the liquid state to the gas state is called vapourisation.

Enthalpy change of vaporisation

The energy required to change one mole of liquid to one mole of gas is called enthalpy change of vapourisation.

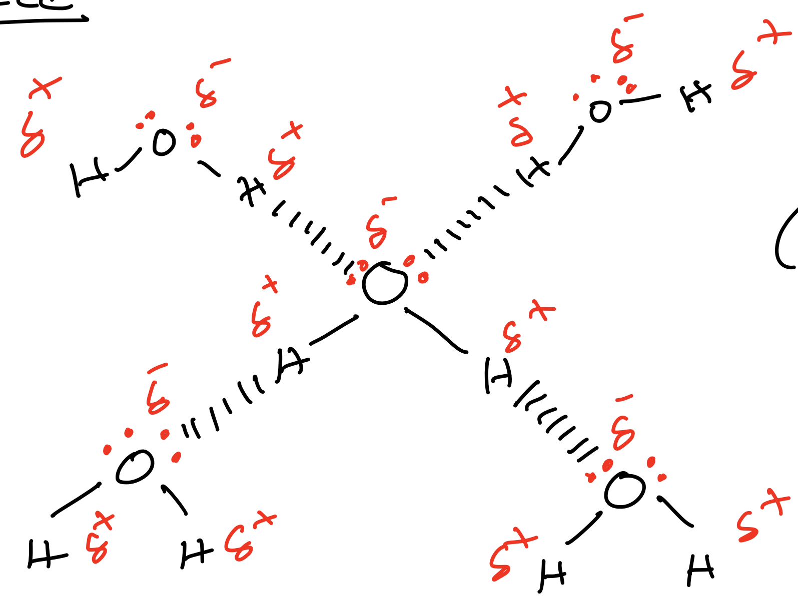
Vapour Pressure

water molecule in liquid $\rightleftharpoons$ water molecules in vapour

- At equilibrium the concentration of water molecules in the vapour remains constant.
- The pressure exerted by a vapour in equilibrium with it's liquid is called vapour pressure.

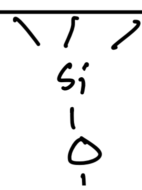
- The vapour pressure is caused by the gas particles hitting the walls of the container.

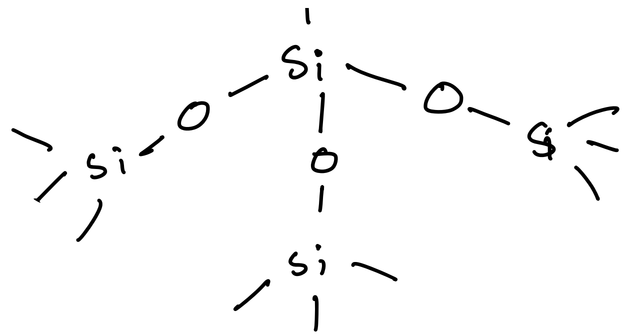
Ice



(simple molecular)

Silicon (IV) oxide / silicon dioxide / silica / SiO_2





- Each silicon atom makes four covalent bonds with four oxygen atoms.

- Each oxygen atom make two covalent bonds with two silicon atoms.

- No delocalised electrons
- Non conductor

- <u>High mp and bp</u>

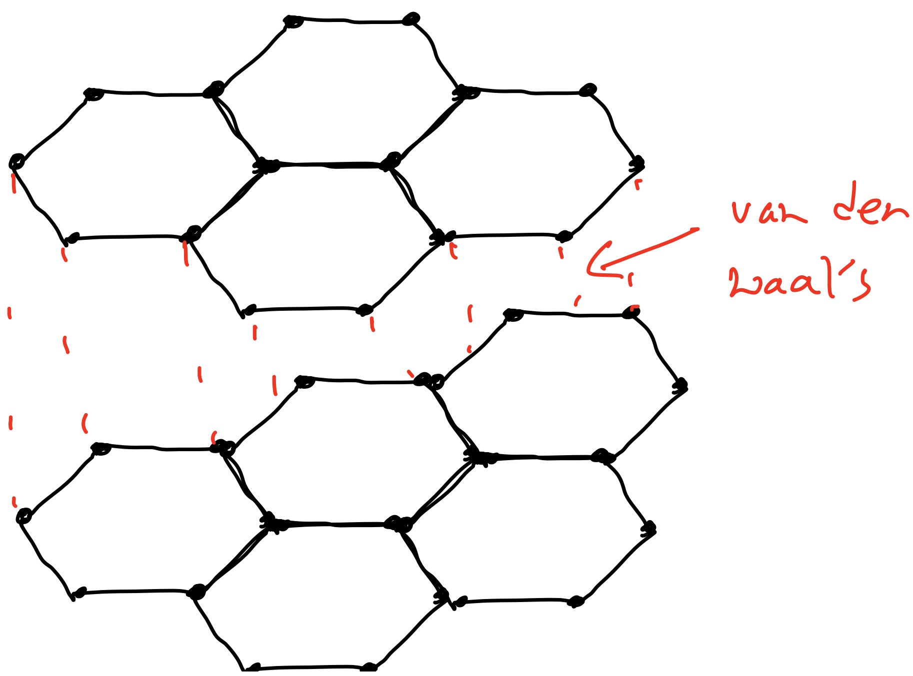
  - Giant covalent structure
  - Many strong covalent bonds.
  - High heat energy is needed to break the strong covalent bonds.

- Insoluble in water.

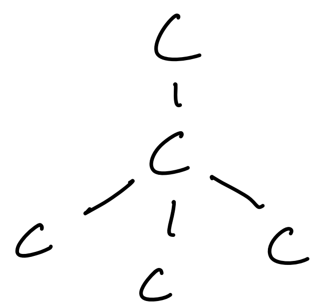
<u>Graphite</u>

- Each carbon atom make three covalent bonds with three carbon atoms.
- Graphite has layer structure
- Layers are made by using hexagons.
- weak van der waal's force of attraction between two layers.
- Layers can slide over each other when force is applied.
- Conduct electricity due to delocalised electrons.

- High m.p
- Carbon atoms are sp^2 hybridized.
- Bond angle 120°.



Diamond



- Each carbon atom is bonded with four carbon atoms.
- Atoms are tetrahedrally arranged
- Carbon atoms are sp^3 hybridized
- Bond angle is 109.5°
- Non conductor of electricity.
- No free electrons / No delocalised electrons
- High mp
  - Giant molecular structure
  - Many strong covalent bonds
  - High energy is needed to break the strong covalent bonds.