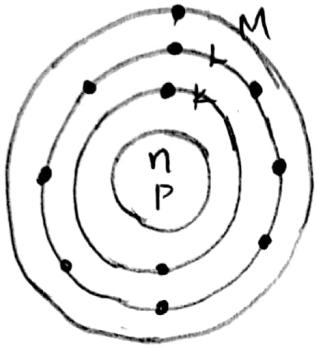
2/12/19

ATOMIC STRUCTURE II & CHEMICAL BONDING

Bohr's model of atom - Postulates.



1. Electrones revolve around nucleus in certain selected circular path called orbits. The energy of an orbit is fixed. ∴ orbit is also known as energy level or quantum level. They are indicated by 1, 2, 3, 4 .... corresponding to K, L, M, N ....

2. As long as electron remains in a particular orbit it doesn't louse or gain energy. ∴ orbit is known as stationary state.

3. Only those orbits are permitted where angular momentum of electron is an integral multiple of $h/2\pi$.

$$mvr = n\left(\frac{h}{2\pi}\right) \qquad n = 1, 2, 3 \ldots$$

3/12/19 4) When an electron jump from one orbit to another it do so by absorbion or emission of energy as radiation, which give spectral lines, whose frequency depends on energy of initial and final levels.

$$h\nu = E_2 - E_1$$

$$\nu = \frac{E_2 - E_1}{h}$$

Merits of Bohrs model

1. It could explain stability of atom. (point 2)
2. It could explain line spectrum or atomic spectrum. (point 4)
3. It help to calculate the radius of n^{th} orbit (point 3)
4. It could ~~explain~~ help to calculate energy of electron in n^{th} orbit.
5. ~~Help~~ Help to explain many lines in H-spectrum.

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De-merits of Bohr's model. (limitations, drawbacks)

1. Inability to explain multi electron spectrum.
2. Inability to explain fine spectrum.
3. Fixed orbit concept is against wave nature and uncertainity principle.
4. Inability to explain bonding and molecular structure.
5. Inability to explain Zeeman effect & Stark effect
6. Inability to explain 3D structure of atom.

Quantum Mechanical model.

i) Dual nature of matter (de Broglie)

According to De-Broglie all material particles in motion behave as particle as well as wave, It is known as dual nature of matter.

Debroglie relation (matter wave equation)

$E = mc^2$ —— (1)

$E = h\nu$ —— (3) $\quad \nu = c/\lambda$

$= \frac{hc}{\lambda}$ —— (2)

$mc^2 = hc/\lambda$

$$mc = \frac{h}{\lambda}$$

$$\lambda = \frac{h}{mc}$$

$\lambda = \frac{h}{mv}$ → This eqn is known as de-broglie eqn.

$\lambda = \frac{h}{p}$

wave length is inversly proportional to mass.

For electron, mass is very small, λ is large wave characters is expected

For ordinary particles, mass is large, λ is very ~~large~~. small wave character is negligible.

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Heisenberg's Uncertainty Principle.

It is impossible to determine both the position and momentum of a sub atomic particle like electron ~~at~~ simultaneously and precisely

$$\Delta x \cdot \Delta P \geq \frac{h}{4\pi}$$

Δx → Uncertainity in position

ΔP → Uncertainity in momentum.

h → planks constant.

$$\Delta x \cdot \Delta(mv) \geq h/4\pi$$

$$\Delta x \cdot m \cdot \Delta v \geq \frac{h}{4\pi}$$

$$\Delta V = \frac{h}{4\pi m \Delta x}$$

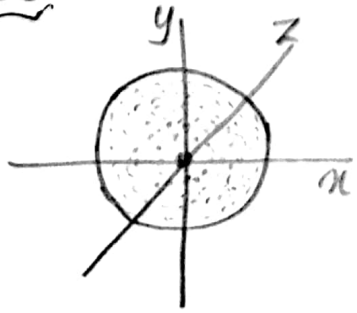
Significans of Uncertainity principle.

The position and velocity of electron cannot be determind simultaneously and correctly. Therefore the fixed orbit is to be replaced by a concept orbital based on probability.

Quantum mechanical model of atom. (orbital)

According to Dual nature, eln poces wave nature, and according to uncertainty principle the position and velocity of eln cannot be determine simultaneously. ∴ The fixed orbit is replaced by a concept 'orbital'.

Orbital:



Orbital is defined as region of space around the nucleus where there is maximum probability of finding an electron.

Difference btw Orbit & Orbital.

Orbit	Orbital.
1. Fixed circular path of electron.	region of space around the nucleus where maximum probability of finding eln.
2. Two dimensional motion of electron	Three dimensional motion of electron.
3. Circular in shape.	different shape (spherical, dumbbell)
4. Maximum no. of eln $= 2n^2$	maximum no. of eln = 2.
5. Against wave nature & Uncertainity principle.	According to wave nature and uncertainity princip
6. could not explain bonding and molecular structure.	can Explain bonding and molecular structur

Quantum Numbers

The electron orbitals are designated by a set of numbers known as quantum numbers. It gives idea about energy, size, shape, orientation and spin of electron orbitals.

1. Principle quantum number. (n)

It indicates the main energy level in which electron present.

$n = 1, 2, 3 \ldots$ (corresponding to K, L, M, N)

It give idea about energy and size of electron orbitals.

2. Subsidiary quantum number.

It indicates the sublevels in a main level

l can have values from zero, up to $(n-1)$. $l = 0, \ldots (n-1)$

$n = 1$ (K)	$n = 2$ (L)	$n = 3$ (M)
$l = 0$, (1s)	$l = 0, 1$	$l = 0, 1, 2$
	2s, 2p	3s, 3p, 3d.

no. of values of l = no. of sublevels.

It give idea about shape of the electron

magnetic quantum number.

It indicates different orientation of a sublevel or orbitals.

$m = -l, \dots 0, \dots +l$ (Total $(2l+1)$ values)

$l = 0$ (s)	$l = 1$ (P)	$l = 2$ (d)
$m = 0$	$m = -1, 0, +1$	$m = -2, -1, 0, 1, 2$
	P_x, P_y, P_z	

4) Spin Quantum number (s)

Electron spin either in clockwise or anticlockwise directions. They are indicated by $+\frac{1}{2}$ and $-\frac{1}{2}$.

Ideas from Quantum number

→ no. of orbitals in a main level $= n^2$

→ no. of eln in a main level $= 2n^2$

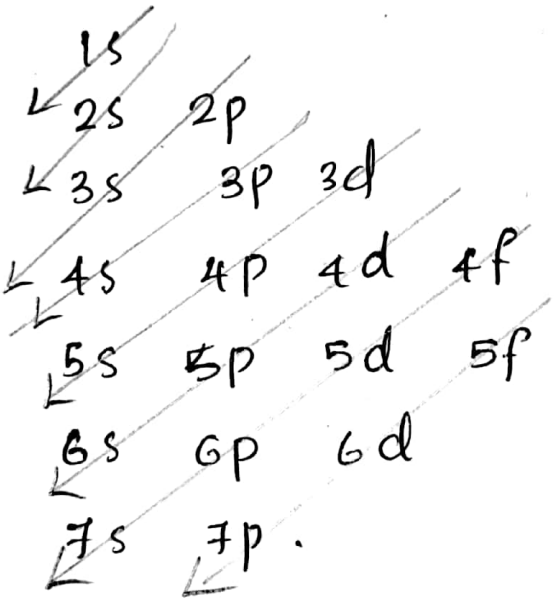
→ no. of orbitals in a sublevel $= 2(l+1)$

→ no. of eln in a sublevel $= 2(2l+1)$
$= 4l+2.$

Electronic configuration.

The distribution of elns in the orbitals according to Aufbau principle, Hund's rule and Pauli's exclusion principle is known as electronic configuration

(1) Afbau principle (building up)

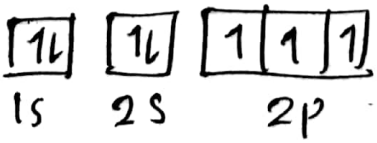


Afbau principle says that elns are filled in the orbitals in the ↑ing order of their energies.

(2) Hund's rule of maximum multiplicity.

No pairing in an orbital of a particular sublevel occur until all available orbitals are half filled.

eg. $_7N$



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Pauli's exclusion principle

It state that no two electrones can have same set of quantum numbers. It means that maximum no. of electrones in an orbital is 2.

eg:- $_1H$ → 1 (1s)

$_2He$ → ↑↓ (1s)

$_3Li$ → ↑↓ (1s) 1 (2s)

$_4Be$ → ↑↓ (1s) ↑↓ (2s)

$_5B$ → ↑↓ (1s) ↑↓ (2s) 1 | | (2p)

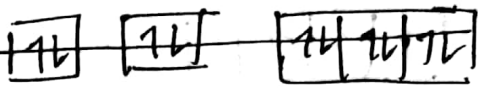
$_6C$ → ↑↓ (1s) ↑↓ (2s) 1 1 | (2p)

$_7N$ → ↑↓ (1s) ↑↓ (2s) 1 1 1 (2p)

$_8O$ → ↑↓ (1s) ↑↓ (2s) ↑↓ 1 1 (2p)

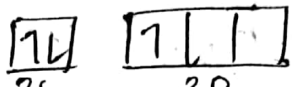
$_9F$ → ↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ 1 (2p)

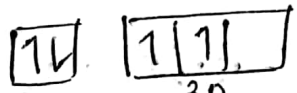
$_{10}Ne$ → ↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p).

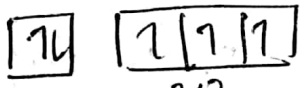
$_{11}Na$ → 

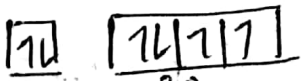
$_{11}Na \rightarrow [Ne]$ 

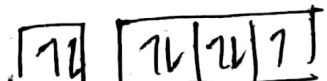
$_{12}Mg \rightarrow [Ne]$ 

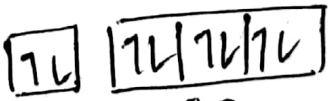
$_{13}Al \rightarrow [Ne]$ 

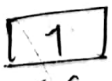
$_{14}Si \rightarrow [Ne]$ 

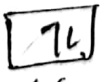
$_{15}P \rightarrow [Ne]$ 

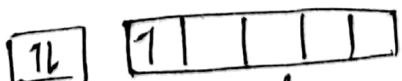
$_{16}S \rightarrow [Ne]$ 

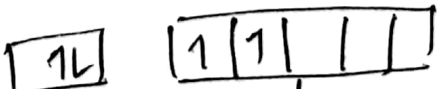
$_{17}Cl \rightarrow [Ne]$ 

$_{18}Ar \rightarrow [Ne]$ 

$_{19}K \rightarrow$ ~~[Ne]~~ $[Ar]$ 

$_{20}Ca \rightarrow [Ar]$ 

$_{21}Sc \rightarrow [Ar]$ 

$_{22}Ti \rightarrow [Ar]$ 

$_{23}V \rightarrow [Ar]$ ↑↓ (4s) ↑ | ↑ | ↑ | | (3d)

$_{24}Cr \rightarrow [Ar]$ ↑ (4s¹) ↑ | ↑ | ↑ | ↑ | ↑ (3d⁵)

$_{25}Mn \rightarrow [Ar]$ ↑↓ ↑ | ↑ | ↑ | ↑ | ↑

$_{26}Fe \rightarrow [Ar]$ ↑↓ ↑↓ | ↑ | ↑ | ↑ | ↑

$_{27}Co \rightarrow [Ar]$ ↑↓ ↑↓ | ↑↓ | ↑↓ | ↑↓

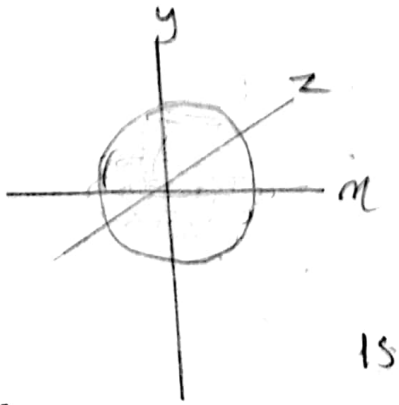
$_{28}Ni \rightarrow [Ar]$ ↑↓ ↑↓ | ↑↓ | ↑ | ↑

$_{29}Cu \rightarrow [Ar]$ ↑ ↑↓ | ↑↓ | ↑↓ | ↑↓

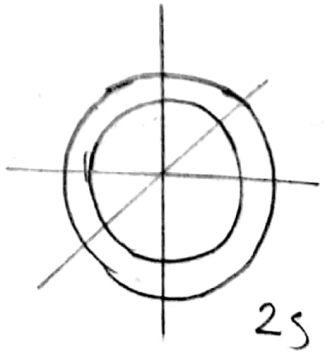
$_{30}Zn \rightarrow [Ar]$ ↑↓ ↑↓ | ↑↓ | ↑↓ | ↑↓

Shapes of Orbitals.

S-orbital.



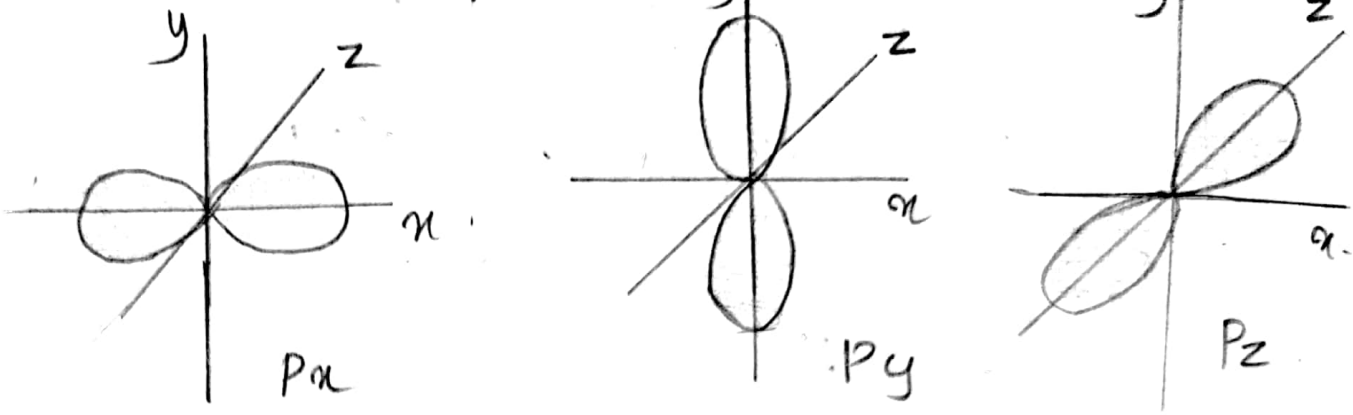
1s



2s

orbitals are spherical shape.. Higher ~~e~~ s orbitals consist of nodes

P-orbital



P orbitals are dumb bell shape.

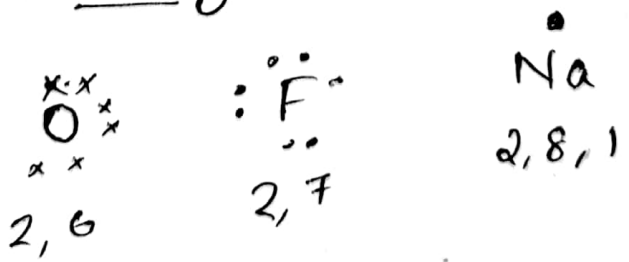
3/1/2020

Chemical Bonding.

It is the force of attraction which bind the constituents (atoms, ions, molecules) together in a substance.

eg:- Ionic bond, covalent bond, coordinate covalent bond, H-bond

Lewis symbols.

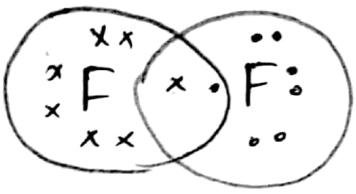


Valency = no. of dots = 8 − no of dots.

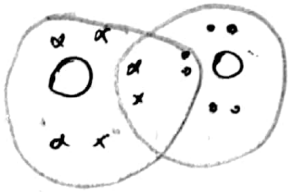
In Lewis symbols the valence electrones are indicated by cross or by dots.

Octet rule. (Lewis & Kossel, electronic theory of valence).

(1) F_2.



(2) O_2.



3) NaCl.

$Na \longrightarrow Na^+ + 1e^-$

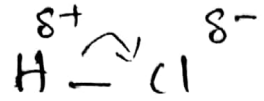
2, 8, 1 2, 8

$Cl + 1e^- \longrightarrow Cl^-$

2, 8, 7 2, 8, 8.

~~Electron~~ atoms take part in chemical bonding to attain 8 elns (octet) in the valence shell. They attain octet either by sharing of electrons or by complete transference of electron.

Electronegetivity.



electronegetivity is the tendency of an atom in a molecule to attract the shared pair of electrones towards itself

- Flurin is the most electronegetive element
- Electronegetivity ries from left → right in a period up to halogens, It dies down the a group.

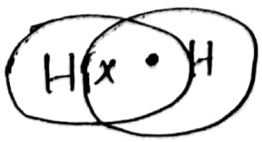
→ increases.

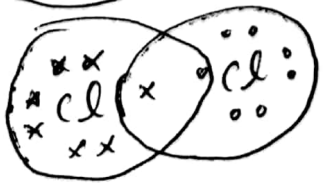
decreases ↓

Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9						
K 0.8						
Rb 0.8						
Cs 0.7						

7/1/20.

Covalent bonding

H_2 →  → H - H

Cl_2 →  → Cl - Cl

The bond formed by mutual sharing of electron pairs between bonded atoms is known as covalent bond.

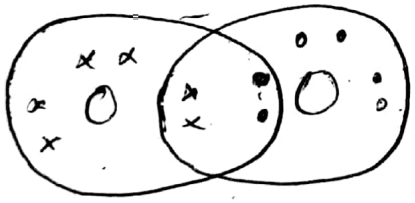
eg:- H_2, Cl_2, O_2, H_2O ... etc.

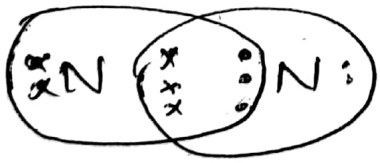
covalent bond is formed between atoms having almost identical electronegetivity values.

multiple covalent bonds:

Double bonds and triple bonds are known as multiple covalent bonds.

eg: O_2 & N_2.

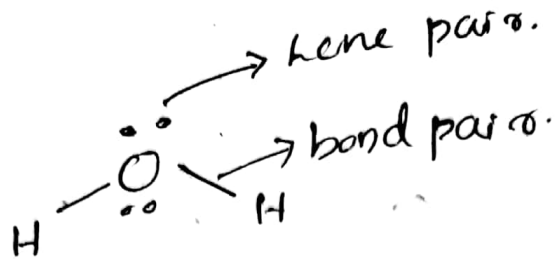
O_2-  $\rightarrow O=O$

N_2.  $\rightarrow N\equiv N$

The bond formed by two electron pairs and 3 electron pairs are known as double bond and triple bond

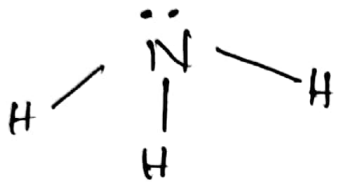
Bond pairs & lone pairs.

1) H_2O



2 bp & 2 lp.

2) NH_3



3bp & 1 lp.

The shared pair of electron is known as bond pair and the unshared pair of eln is known as lone pair.

Ionic Bonding

eg: 1) NaCl.

$$Na \longrightarrow Na^+ + 1e^-$$

2,8,1

$$Cl + 1e^- \longrightarrow Cl^-$$

2,8,7 — 2,8,8.

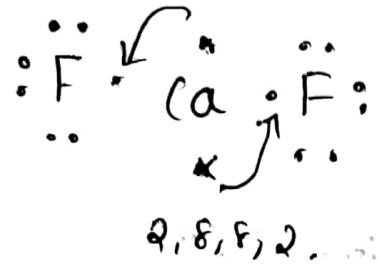
$[Na^+][Cl^-]$

oppositely charged ions are formed by complete transference of electron. The electrostatic force of

attraction between these oppositely charged ion is known as ionic bond.

eg: CaF_2, MgO

CaF_2.



$Ca \longrightarrow Ca^{2+} + 2e^-$

$2F + 1e^- \longrightarrow 2F^-$

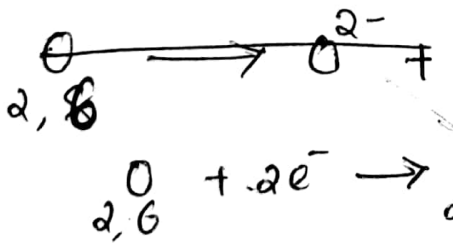
2, 8, 8, 2.

$[F^-][Ca^{2+}][F^-]$.

MgO.

$Mg \longrightarrow Mg^{2+} + 2e^-$

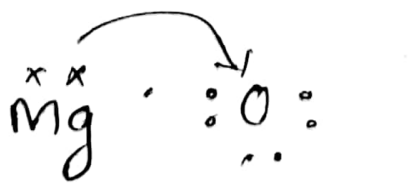
2, 8, 2



2, 8

$O + 2e^- \longrightarrow O^{2-}$

2, 6 2, 8.



2, 8, 2. 2, 6.

$[Mg^{2+}][O^{2-}]$

Expiry date

Electro Valency

The no. of electrons lost or gained by an atom in ionic bonding is known as electro valency.

eg: Na - 1 , Mg - 2 , Cl - 1 , O - 2

10/1/20

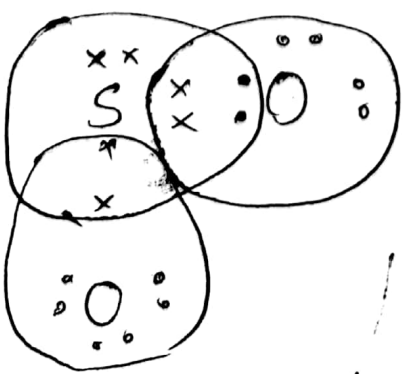
Conditions for ionic bond formation.

1. metal atom should be highly electropositive (low ionization enthalpy for metal atom.)
2. Non metal atom should be highly electronegetive. (high -ve eln gain enthalpy)
3. high -ve lattice enthalpy. (amount of energy released when 1 mol of solid ionic crystal is formed from its gaseous ions.

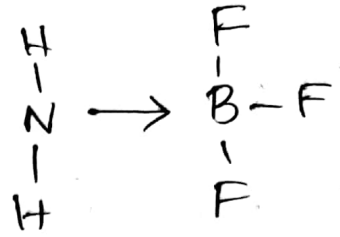
Ionic compounds are formed btw elements of group 1 or 2 with elements of group 16 or 17

Co-ordinate bond.

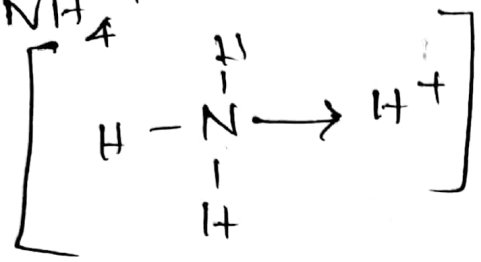
1) SO_2



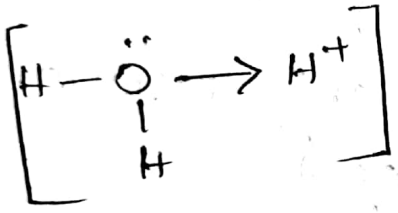
S = O
↓
O

2) $H_3N + BF_3 \longrightarrow$ 

3) NH_4^+



4) H_3O^+

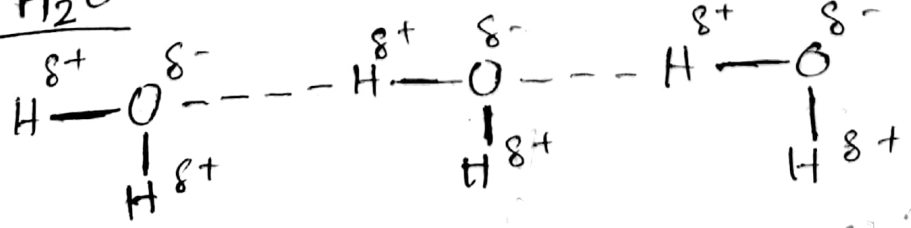


→ The bond formed by sharing of electron pair which is contributed by one of the atom in the molecule is known as co-ordinate bond.

→ It is indicated by an arrow, starting from donor ending at accepter.

Hydrogen bonding.

H_2O



The bond formed by H-atom of one molecule with electronegetive (F, O, N) atom of another molecule is known as H-bond.

eg:- H_2O, HF, NH_3

Consequencies of H-bonding.

1. The liquid state of water compared to gaseous state of H_2S is due to association by intermolecular H-bonding.
2. The higher boiling point of alcohol compared to ether is due to association by intermolecular H-bonding.

Factors Affecting magnitude of H-bonding.

1) Electronegetivity of the element

$F > O > N$.

2) Size of the electronegetive atom.

~~Sta~~ smaller the size stronger the H-bonding. (Nitrogen form H-bond whereas Cl doesn't)

3) states of the substance.

Solid > liquid > gas.

H-bonding is strong & in solid state.

Effects & Properties of H-bonding

1. Water in plants and animals is attached to proteins by H-bonding.
2. The structure of proteins and nucleic acids is stabilized by H-bonding.