# Atomic Structure and Chemical Bonding.

Syllabus →

1. structure of atom mass Number and atomic Number, Isotopes and octate Rule.

- Definition of an atom
- Constituents of an atom – Nucleus (Proton, Neutron)
- Electron Distribution in the orbits
- Definition and Examples of Isotopes (H, C, Cl)

1. Introduction →

Many attempts have been made to know as to what is the ultimate particle of matter.

→ The Idea of smallest unit of matter was first given by Maharshi Kanada in 6th century B.C. in India. a/c to him matter consisted of Indestructible minute particles called 'Parmanus', now called atom

→ The Greek philosopher Democritus (460 B.C. – 370 B.C) called the Parmanu an "atom"

→ the first scientific theory about the structure of matter was given by John Dalton (in 1808)
It considered atom as Indivisible particles that are ~~the structure of matter (atom).~~

the fundamental building blocks of matter.

the main Postulates of Dalton's theory are-

1. matter consist of very small & indivisible

2. atom of an element alike in all Respects but they Differ from of other elements.

3. atom of an Elements Combine in small numbers to form molecules.

4. Atom are the smallest units of matter that can take part in a chemical Reaction.

→ atoms consists of subatomic Particles i.e. ① Electron ② Protone ③ Neutrons.

| Sub atomic Particles. | charge | Discovered by. |
| --- | --- | --- |
| 1. Electron | -ve | J.J. Thomson |
| 2. Protone | +ve | E. Goldstein |
| 3. Neutron | Neutral | James Chadwick |

⊛ Definition of an element →

An element is a substance which is made up of only one type of atoms.

For Example- C is an element bz it cannot be split up into 2 & more simpler substance

Exception → Radioactivity

★ Definition of atom →

" an atom is the smallest particle of an element that exhibits all the properties of that element. It may or may not exist Independently but takes part in every chemical rxn.

Exa → Take a small piece of zinc and grind into smaller pieces. all these pieces show properties of zinc. But on grinding these further they break up into very fine particles which still show the properties of zinc. But there came a stage when the particles cannot be further subdived into particles exhibiting properties of zinc.

⊛ Constituents of an atom →

Discovery of electrons →

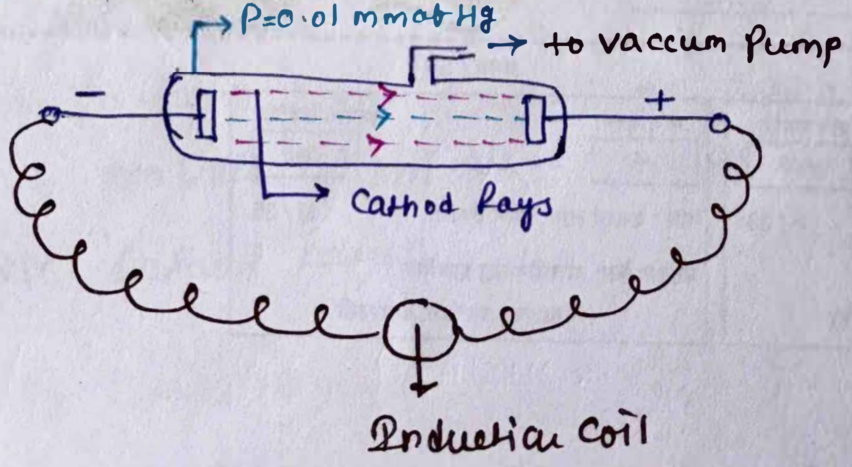



Fig.01

First Studied about Electron by william Crooke, Later J.J. thomsons Studied the characteristics and the Constituents of cathod Rays.

The apparatus used by Him is called "Discharge tube" or a Cathod ray tube shown in fig. 01

a Discharge tube is a hard glass fitted with two metals plates as electrodes, one of which connected to the positive terminal of Battery and is called Anode (⊕ve), while the other electrode which is connected to the negative terminal of the Battery is called Cathod (⊖ve), It has a side tube through which gas can be Pumped out by using a Vacuum pump to create Vacuum.

Scientist filled a gas in the vacuum tube $\bar{c}$ High Voltage (10000 v) & Very low Pressure (0.01 mmHg), gas are ordinarily Poor Conducter of electricity But at High Voltage It become a good Conducter of electricity and begin to flow from cathod to anode in the form of Rays & travels from the cathod to anode they are called "Cathod Rays"

* Properties of Cathod Rays-

→ they travel from the cathod to the anode in straight lines.

They are affected by electric field i.e. they are attracted towards the positive field and deflected from the negative field. Cathode rays shows that they carry -ve charge.

→ Cathode rays is made to fall upon hard metallic targets like W (tungsten), X-rays are produced.

# Thomson's concluded that -

1. Cathode rays consist of -ve charged particles now called $e^-$s (electrons)

2. They are an integral part of all atoms.

3. $e^-$ have both definite mass and definite electric charge

# Properties of electrons →

→ The mass of $e^-$ ⇒ 1/1837 the mass of a hydrogen atom ($9.18 \times 10^{-31}$ kg)

→ charge of $e^-$ → $1.602 \times 10^{-19}$ C

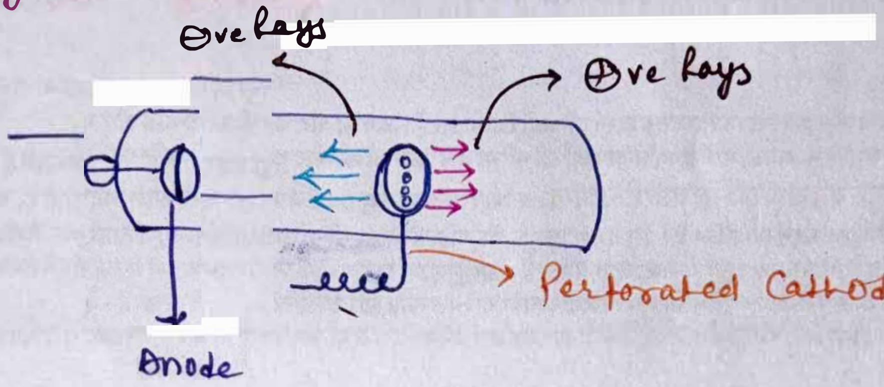
→ Radius is less than $1 \times 10^{-15}$ m.

Atoms are found to be electrically neutral, so they must contain, in addition, particles that are +ve charged, such that the -ve charge of the $e^-$s is equal to the total +ve charge.

This realization led to the discovery of +ve charged subatomic particles protons.

## * Discovery of Protons →

Goldstein noticed that another set of rays travelling in a direction opposite to that of the cathode rays i.e. from anode (⊕ve) toward cathode, when a perforated cathode was used in the above discharge tube. He called these rays as Canal Rays/⊕ve Rays/ anode Rays.




## * Properties of Anode Rays →

→ Anode Rays made up Positively Charged Particles.

→ ⊕ve Rays are deflected by electric and magnetic field but in a direction opposite to that of the Cathode Rays. this means that these rays consist of ⊕ve charged Particles called Protons.

→ These are produce fluorescence on a zinc sulphide screen

→ Their e/m i.e charge to mass Ratio, differs from gas to gas, Its value is much less than that of an e- & is maximum when H is taken in the Discharge tube

* Properties of Protons →

→ A Proton possesses a unit of ⊕ve charge +1 the value of $(1.602 \times 10^{-19})$ C

→ mass ⇒ $(1.672 \times 10^{-24} g)$

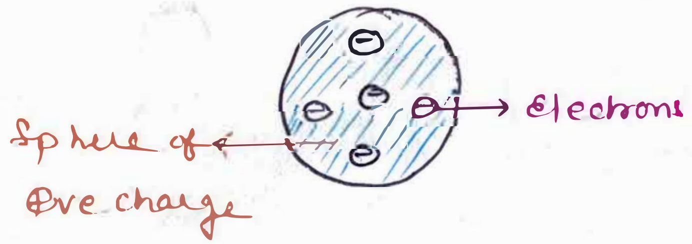
# After the Discovery of e⁻s and Protons J.J. thomson Proposed 'plum Pudding' model of the atom.

a/c to this model →

1. An atom is Considered to be a sphere of uniform ⊕ve charge and e⁻s are embeded into it.

2, Total ⊕ve charge = total ⊖ve charge that an atom as a whole is Electrically neutral.

3. The mass of an atom is Considered to be uniformly Distributed.




(*) Discovery of Nucleus →

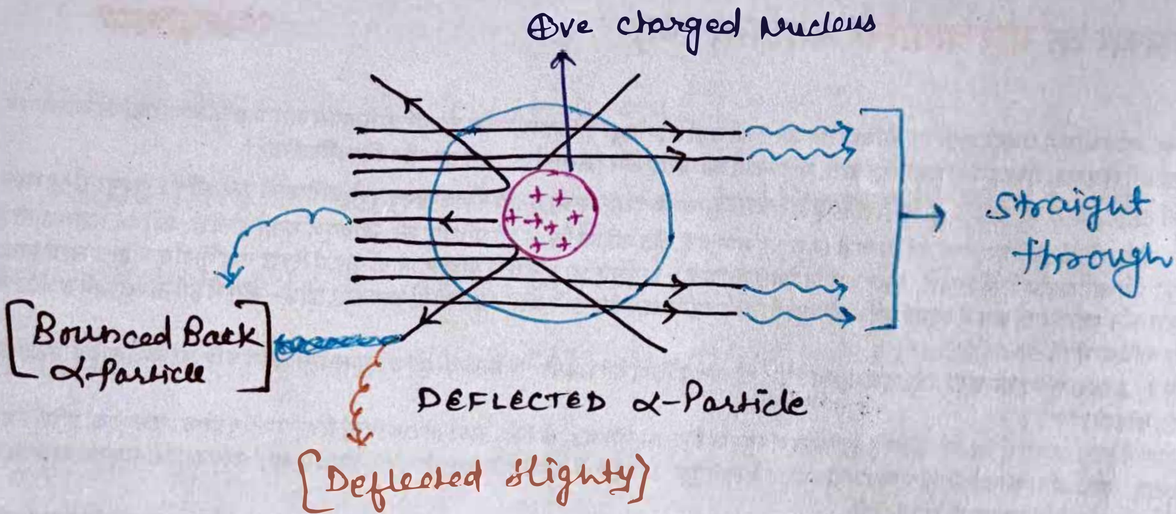
In 1911, Rutherford, Directed a stream of α-Particles towards a very thin gold foil. He selected a gold foil Bz he wanted a thin a layer as Possible and gold is the most mellable metal.

He observed that-

① Most of the α-Particles passed straight through the foil.

② Some α- Particles were slightly deflected from their Straight path

③ Very few α-Particles were either deflected by very Large angles or completely bounced back.




these Result of α-Particles Scattering Experiment and suggested a model of atom that is k/a Rutherford's Atomic model.

# a/c to this model:-

→ The atom contain a large empty space. this is why most of the α- Particle through the metal foil without deviating from their path

→ There is a ⊕ve charged mass at the centre of atom k/a nucleus. In which the entire mass of the atom is concentrated.

Positively charged α-Particles approaching the nucleus get Deflected.

→ The size of the nucleus is very small compared to the size of an atom so the α-Particles pass straight through the gold foil.

→ e⁻s Revolve around the nucleus in close circular Path (orbits). the force of attraction b/w the -ve charged e⁻s and the Positively charged nucleus is counter balanced by the centrifugal force

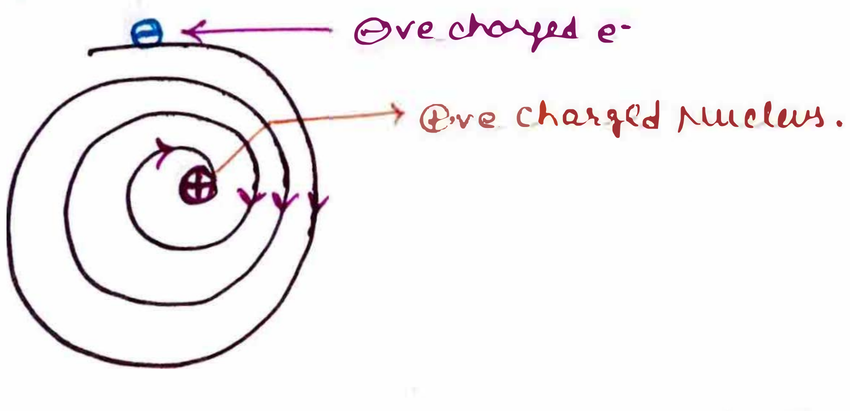
→ Rutherford's model of atomic structure is similar to the structure of the solar system, the sun is at the centre and the planets revolve around it

* Drawback of Rutherford's atomic model →

1.→ Comparison of e⁻s ē planets in the solar system

2.→ an e⁻, when moving around the nucleus continually should radiate energy i.e. loses energy

It sh. be gradually pulled towards the nucleus and end up colliding it. this should result in the total collapse of the atom.



If it was so, the atom should be highly unstable and 
Hence matter would not exist in the form.

Thus Rutherford's model failed to explain the stability of an atom.

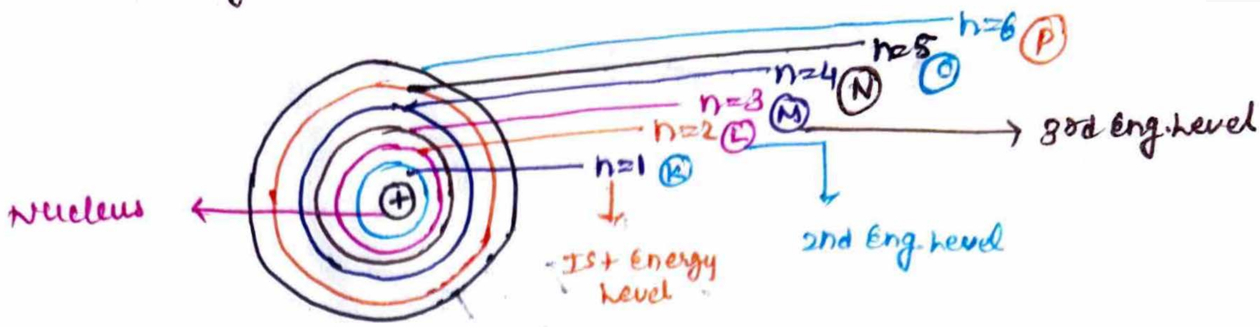
## ⊛ Bohr's Atomic model →

In 1913, Niels bohr, explained the causes of the stability of the atom in a different manner.

**Postulates →**

1. The e⁻ revolving around the nucleus are confined to certain orbits called shells.
2. While it is revolving around the orbit, e⁻ neither lose nor gains energy.
3. An e⁻ revolving in a particular orbit, on gaining a certain amount of energy, jumps to the next orbit and vice versa.
4. Each orbit is associated $\bar{c}$ a fixed amount of energy. Bohr called it an energy level.
5. These energy level are labelled K, L, M, N or. I, II, III, IV etc., the orbit closest to the nucleus is the K-shell.
6. It has least amount of energy and the e⁻s present in it are called K e⁻s., and so on $\bar{c}$ the successive shells and their e⁻s.

→ Energy Levels around the Nucleus of an atom :-




## (*) Discovery of Neutron →

In 1932, Chadwick discovered these particles by bombarding light nuclei like Be $\bar{c}$ $\alpha$-particles i.e. He nuclei

$${}_4Be^9 + {}_2He^4(\alpha) \longrightarrow {}_6C^{12} + {}_0n^1 \text{ (Neutron)}$$

These particles are found to be neutral, so named Neutrons.

### * Properties of Neutrons →

→ This particle was not found to be deflected by any magnetic or electric field, proving that it is electrically neutral.

→ Mass ⇒ $1.676 \times 10^{-24}$ g (1 amu)

→ Neutron are electrically neutral particles that are also found in nucleus.

→ Neutron are slightly heavier than protons.

## * Atom - Its Structure -

→ The atom is built up of a number of sub-atomic particles.

→ The 3 sub-atomic particles of great importance in understanding the structure of atom are $e^-$, $p^+$, $n^0$

→ There are two structural part of an atom, the nucleus and the empty space in which there are imaginary paths called orbits.

→ Nucleus → Protons + neutrons ⇒ Nucleons are found in the central part of atom

### ⊛ Atomic Number (Z) →

→ Atomic Number (Z) = No. of Protons ($p^+$)

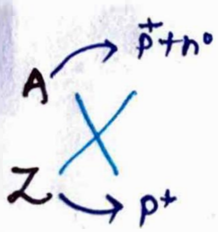
→ The (Z) is, therefore, equal to the No. of e− in the atom of an element as well, since in a Neutral atom

No. of Protons = No. of $e^-$s

* Symbolic Representation of element

[Z]

[A] mass No.




atomic No. [Z]
= No. of Protons [p]
= No. of electrons [e]

$Z = P = e$

⇒ No. of [$n^0$] + No. of [$p^+$]

⇒ $A = Z + P$

Example → $^{35}_{17}Cl$ →

$Z ⇒ p^+ ⇒ e^- ⇒ 17$

$A ⇒ n + p ⇒ 35$

$n^0 ⇒ A - Z ⇒ 18$

* Mass Number [A] - Total No. of $p^+$ & $n^0$ present in its nucleus

* Distribution of $e^-$s in the orbits - Bohr-bury Scheme →

Distribution of $e^-$s into Different shells (orbits) of an atom was suggested by Bohr-bury.

# The following Rules are followed →

1. The maximum Possible Number of $e^-$s in a Particular shell is given by the formula - $2n^2$

$n$ ⇒ shell No.

Ist or K shell → $(2 \times 1^2) = 2$ $e^-$s

2nd or L shell → $(2 \times 2^2) = 8$ $e^-$s

3rd or M shell → $(2 \times 3^2) = 18$ $e^-$s

4th or N shell → $(2 \times 4^2) = 32$ $e^-$s

Exa → Ca (calcium) ⇒ Having 20 $e^-$s

| K | L | M | N |
|---|---|---|---|
| 2 | 8 | 8 | 2 |

* Structure of mg (magnisium ion) →

$^{24}_{12}Mg$ →

No. of $p^+$ ⇒ 12 (equal to Z) = $e^-$s

No. of $p^+ + n^0$ ⇒ 24 (Equal to A)

No. of $n^0$ ⇒ $(24-12) = 12$

It is surrounded by 12 $e^-$s that are allowed to Different shells as follow:

K shell → 2 e⁻s

L shell → 8 e⁻s , Electronic Configuration ⇒ 2, 8, 2

M shell → 2 e⁻s

→ In lighter elements, up to Ar [Argon] [Z = 18], Each Inner Shell is completely filled before any e⁻ can occupy an outer shell.

However, In elements heavier than Argon (Ar), the situation changes.

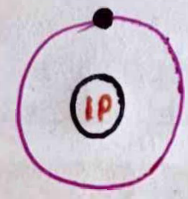
although the 3rd shell can accomodate up to 18 e⁻s yet the 4th shell begins to be filled after it has only 8 e⁻s

Eg → K = (Z = 19)

| K | L | M | N |
|---|---|---|---|
| 2 | 8 | 8 | 1 |

* Some orbital Diagrams →

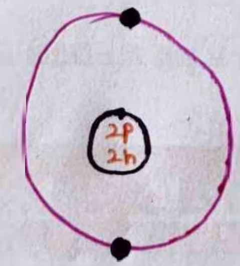
Atom of $^{1}_{1}H$




Electronic Config$^{n}$ ⇒ 1

Valency ⇒ 1

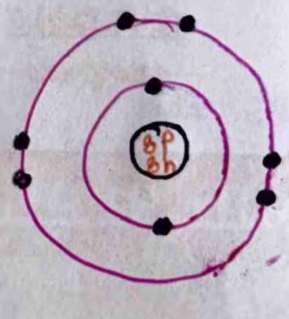
atom of Helium $^{4}_{2}He$




E.C ⇒ 2

V ⇒ 0

Atom of oxygen $^{16}_{8}O$




E.C ⇒ 2, 6

V ⇒ −2

# ⊛ Valence Electrons →

The outermost shell of an atom is called its valence shell, and e⁻s ~~out~~ in the valence shell is called valence e⁻s.

→ H has only 1e⁻ in it valency shell
→ C " ———" 4e⁻ " ———" ———"
→ Chemical properties of elements are decided by their valence e⁻s, since they are the ones that take part in chemical reaction.

→ Elements can lose e⁻ to form ions which are positively charged are called cation. (+)

Eg → $Na - e^- \longrightarrow Na^+$
2,8,1 → 2,8

→ Those elements can gain e⁻s to form ion which are negatively charged and are called ANIONS (−)

Eg → $Cl + e^- \longrightarrow Cl^{\ominus}$
2,8,7 → 2,8,8

They are good oxidising agents

→ elements having 8e⁻ in valence shell called Inert gases which is inactive chemically.

## ⊛ Reason of chemical activity of an atom →

Atoms of all noble gases (except-He) have 8 e⁻s in their outermost shell. This arrangement is k/a OCTET.

| Inert gas | [Z] | Electronic Configuration K | L | M | N | O | P |
|---|---|---|---|---|---|---|---|
| He | 2 | 2 | | | | | |
| Ne | 10 | 2 | 8 | | | | |
| Ar | 18 | 2 | 8 | 8 | | | |
| Kr | 36 | 2 | 8 | 18 | 8 | | |
| Xe | 54 | 2 | 8 | 18 | 18 | 8 | |
| Rn | 86 | 2 | 8 | 18 | 32 | 18 | 8 |

→ They donot combine c̄ other atom to form molecules.

→ Their valency is zero.

→ Chemically active atoms have an incomplete octate.

→ The atom of all elements, other than the inert gases, combine bz they have incomplete valence shells and tend to attain a stable configuration.

→ The number of e⁻s gained, lost or shared to attain the octate in the outermost shell, gives the combining capacity of the element that is, its valency.

⊛ **Isotopes →** Isotopes may be defined as atoms of the same element having the same atomic number but different mass numbers.

## * Examples of Isotopes →

| Element | NO. of Isotopes | |
|---|---|---|
| 1. Hydrogen | 3 | $^{1}_{1}H$ , $^{2}_{1}H$ , $^{3}_{1}H$ |
| 2. Carbon | 3 | $^{12}_{6}C$ , $^{13}_{6}C$ , $^{14}_{6}C$ |
| 3. Chlorine | 2 | $^{35}_{17}Cl$ , $^{37}_{37}Cl$ |
| 4. Oxygen | 2 | $^{16}_{8}O$ , $^{18}_{8}O$ |
| 5. Potassium | 2 | $^{39}_{19}K$ , $^{41}_{19}K$ |

## * Uses of Isotopes →

→ $^{60}_{27}Co$ → Radiotherapy for treating Cancer and other disease.

→ $^{14}_{6}C$ → Determining the age of Historical & geological Material

→ $^{235}_{92}U$ → Fuel in Nuclear Reactors.

→ $^{131}_{53}I$ → Treatment of Goitre.

* **Isobars →** They are atom of Different Elements with the Same Mass Number (A), but Different atomic (Z) numbers.

Z → Different

A → same

Eg → $^{40}_{18}Ar$ & $^{40}_{20}Ca$ , $^{54}_{24}Cr$ & $^{54}_{26}Fe$

⊛ Chemical Bond → A chemical Bond may be defined as the force of attraction b/w the two atoms which binds them together as a unit called molecule.

→ The chemical Combination of atoms Involves Redistribution of e⁻s so as to leave each atom with a Stable electronic Configuration.

⊕ Electrovalent (OR Ionic) Bond →

The Cation and Ionic (anion) being oppositely charged attract to each other a chemical Bond, Since this chemical Bond formation is due to the electrostatic force of attraction b/w c⊕ (cation) and an anion (⊖ve), It is called Electrovalent (or Ionic) Bond.

* Electrovalency → The No. of e⁻s that an atom of an Element loses or gains to form a electrovalent Bond is called its electrovalency.

* Metals → 1, 2 & 13 Group No. ↪ Tendency to lose their e⁻s.

* Non-Metals → 15, 16 and 17 → Tendency to gain e⁻s.

⇒ Group → 17th ⇒ Most electronegative elements [F]
CsF → Most Ionic Compound

Metals ⟶ Electrovalent Bond ⟵ Non-Metals

Q. why are Ionic Compounds stable?

A→ Electrostatic force of attraction b/w opposite charges is much higher, It makes the Ionic Compounds stable.

* Examples of Electrovalent compounds→

$NaCl$

$MgCl_2$

$CaO$

$KBr$

$CaCl_2$

* Structure of some electrovalent Compounds→

(1) NaCl (Sodium chloride)→

Electronic configuration→ Na ⇒ 2, 8, 1

It has one e- in excess of the stable electronic configuration of the Nearest Nobal gas (Ne) neon (2, 8), Therefore Na shows a tendency to give up the e- from outermost shell.

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

(2, 8, 1) (2, 8)

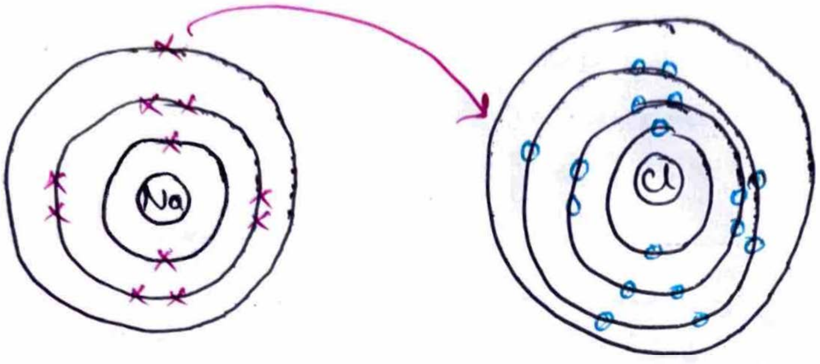
atom cation

• Chlorine (Cl)→

- Electronic Configuration ⇒ 2, 8, 7
- one e- less than that of nearest noble gas, Ar (2, 8, 8)
- Cl shows tendency to [crossed out] aquire one e- to attain octet.

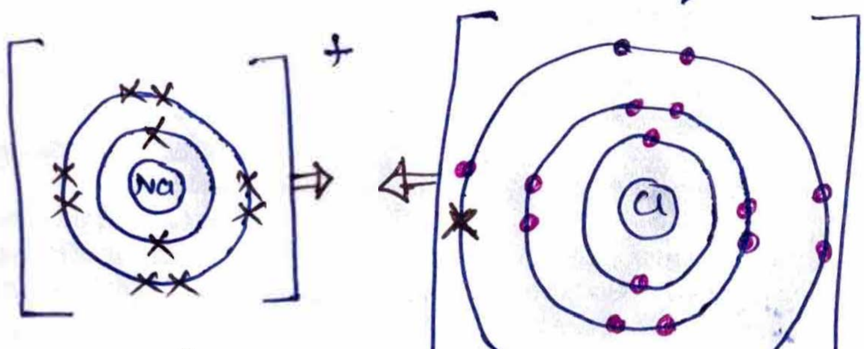
Q→ $Cl + 1e^- \rightarrow Cl^{\ominus}$

(2, 8, 7) atom → (2, 8, 8) anion




$Na \rightarrow Na^+ + e^-$
(2, 8, 1)

$Cl + e^- \rightarrow Cl^{\ominus}$
(2, 8, 7)




$[Na]^+$
(2, 8)

$[Cl^-]$
(2, 8, 8)

[Orbit structure of Electrovalent bond in NaCl]

Cation $[Na^+]$ & Anion $[Cl^-]$ are attracted towards to each other.

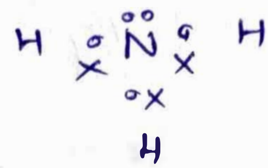
## ⊛ Electron Dot Symbol (Lewis Symbol)→

The $e^-$ Dot symbol for an atom consists of the symbol of the element surrounded by dots representing the outermost shell $e^-$s.

The paired e⁻s are represented by a pairs of dots, whereas the unpaired e⁻ in the outmost orbit is represented by a single dot.

Exa → e⁻ dot symbol of H is $H^{\circ}$ and of oxygen is :O:

Exa → Ammonia ($NH_3$) ⇒



(2) magnesium chloride ($MgCl_2$) →

Mg(Z) → 12

Valence e⁻ of Mg → 2 (●)

Cl (Z) → 17

Valence e⁻ of Cl → 7

Mg atom aquires a stable configuration of 8 e⁻ by losing 2 e⁻s from its outmost shell & thus becomes a +ve Mg ion $Mg^{+2}$

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

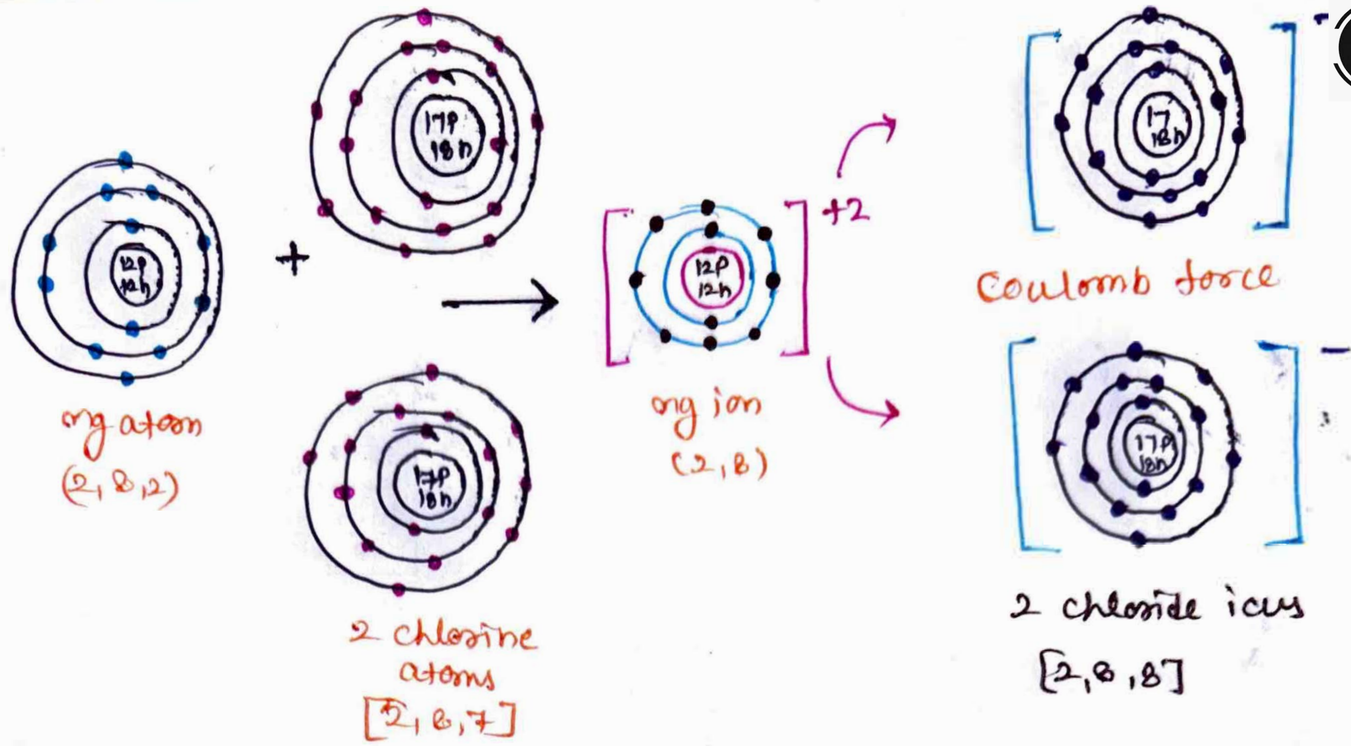
(2, 8, 2) atom → (2, 8) cation

Cl atom, which contains 7 e⁻s in its outermost shell, can accept only 1 of the 2 e⁻s donated by a Mg ~~ion~~ atom

Mg atom forming a Mg ion, there must be 2 Cl atoms to form 2 chloride ions.

$$2Cl + 2e^- \longrightarrow 2Cl^-$$

(2, 8, 7) 2Cl atoms → (2, 8, 8) 2 Cl⁻




→ Mg : Cl

1 : 2

✱ Electron Dot Structure of $MgCl_2$ →

$Mg: + 2\,Cl \longrightarrow Mg^{+2}\ 2[:\ddot{Cl}:]^{-}$

✱ Calcium Oxide (CaO) →

| Ca | O |
|---|---|
| Z = 20 | Z = 8 |
| Valancy = 2 | Valancy = 6 |
| Ca atom loses 2 e-s to attain a octet | Oxygen requires 2 e-s to attain octet |
| $Ca^{+2}$ | $O^{-2}$ |

Only 1 oxygen atom is needed to accept the 2 valence e-s donated by a calcium atom, the formula of calcium oxide is CaO and not $Ca_2O_2$

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

(2, 8, 8, 2) atom → (2, 8, 8) Cation

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

(2, 6) → (2, 8)

* Electron Dot structure of Calcium oxide →

$Ca: + \ddot{O}: \longrightarrow Ca^{+2}\ [:\ddot{O}:]^{-2}$

* In the formation of a electrovalent bond, the transfer of e- is involved.

The Electropositive atom undergoes Oxidation, while the electronegative atom undergoes Reduction, this is a Redox reaction.

Eg → Formation of NaCl →

$2Na + Cl_2 \longrightarrow 2Na^+ + 2Cl^-$ (or 2NaCl)

written as two Half Reactions:-

$2Na \longrightarrow 2Na^+ + 2e^-$ (oxidation)

$Cl_2 + 2e^- \longrightarrow 2Cl^-$ (Reduction)

---

$2Na + Cl_2 \longrightarrow 2Na^+ + 2Cl^-$ [Redox Reaction]

Na → $Na^+$ [O]; $Cl_2$ → $Cl^-$ [R]

Eg → $CuO + H_2 \longrightarrow Cu + H_2O$

$H_2$ → $H_2O$: oxidised

CuO: Oxidising Agent; $H_2$: Reducing Agent

CuO → Cu: Reduced

$Cu^{+2} + 2e^- \longrightarrow Cu$ (Reduction)

$2H - 2e^- \longrightarrow 2H^+$ (oxidation)

# ⊛ Covalent (molecular) Bond →

The chemical bond that is formed b/w two combining atoms by mutual sharing of one or more pairs of e⁻s is called a Covalent Bond. and the compound formed due to this bond is called a Covalent Compound.

The molecule formed due to the sharing of e⁻s (Covalent Bond) is called a Covalent molecule.

The sharing of e⁻s takes place b/w their atoms and a covalent Bond is formed.

eg → (H) is a non-metal and chlorine is also a non-metal when Hydrogen combines c̄ chlorine to form Hydrogen chloride (HCl), due to sharing of e⁻s take places b/w H & Cl atoms. and a covalent bond formed.

⊛ Covalent bonds are of 3 types :-

1. Single Covalent bond
2. Double Covalent bond
3. Triple Covalent bond

1 Single Covalent → sharing of one pair of e⁻s b/w the atoms Each atom contributing one e⁻.

→ Denoted by putting a short line (–), Like H–H

Cl–Cl, H–Cl, H–O–H ($H_2O$), H–N(–H)–H (Ammonia), H–C(–H)(–H)–H (methane)

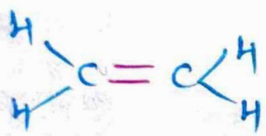
2. Double bond → sharing of 2 pairs of e⁻ b/w 2 atoms.

Eg – $O_2$, $(O=O)$

3. Triple bond – Combination of 3 single bonds.

Eg – $N_2$, $(N\equiv N)$

Some molecule have a combination of single & Double Triple / Both bonds –



ethene ($C_2H_4$)

$H-C\equiv C-H$

ethyne ($C_2H_2$)

⊛ Non-Polar & Polar Covalent Compounds →

If the 2 covalently bonded atoms are identical the shared e⁻ pair is at equal distance from the combining atom i.e. the shared e⁻ pair is equally attracted by the nuclei of the 2 types of charge

Eg → $H_2$, $Cl_2$, $O_2$ etc.

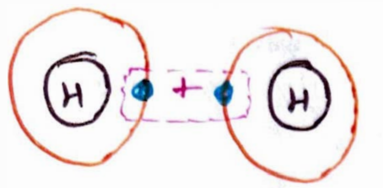
– The covalent compounds are said to be Polar when the shared pair of e⁻s are not at equal distance b/w the 2 atoms.

Eg → $\overset{\delta+}{H}-\overset{\delta-}{Cl}$

## ⊛ Some Covalent molecules and their Structures →

### 1. Hydrogen molecule (Non-Polar compound)

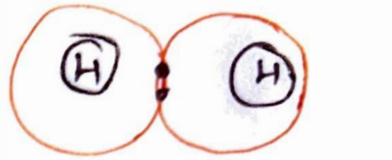
Before Combination




e⁻ Dot Structure
↓

H • + • H

after Combination



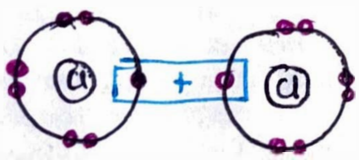

H : H or H — H ⟶ $H_2$

### 2. Chlorine - chlorine (Non-Polar Comp.) →

$_{17}Cl$ ⇒ 2, 8, 7

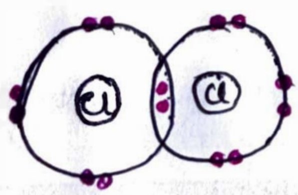
Nearest noble gas → Ar (Argon)
(2, 8, 8)

e⁻ Dot Structure →




:C̈l• + •C̈l:

after Combination –




Chlorine molecule

:C̈l: C̈l: or [Cl — Cl] ⟶ $Cl_2$

Single Covalent Bond

### 3. Nitrogen molecule –

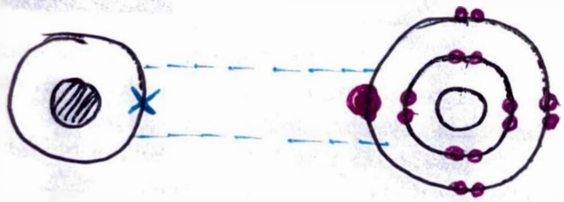
⊛ Hydrogen chloride (H-Cl) → Polar Compound

$[H] \rightarrow z = 1$ Nearest gas (Noble) → Ar [2, 8, 8]

$_{17}Cl \rightarrow [2, 8, 7]$

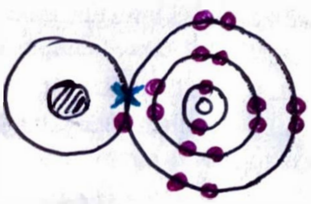
To attain the stable electronic configuration of nearest noble gas hydrogen needs one e- and chlorine also needs one e-.

Before Combination



$H^{\times}$ $\cdot\ddot{Cl}:$

after Combination-



H · ×Cl: or H-Cl

Hydrogen chloride molecule

one molecule of H-Cl contains a total of 2 atom i.e. one atom of H & one atom of chlorine.

Eg → water ($H_2O$) molecule

Eg → Ammonia molecule (Polar compound)

Eg → Carbon tetrachloride (Non-polar comp.)

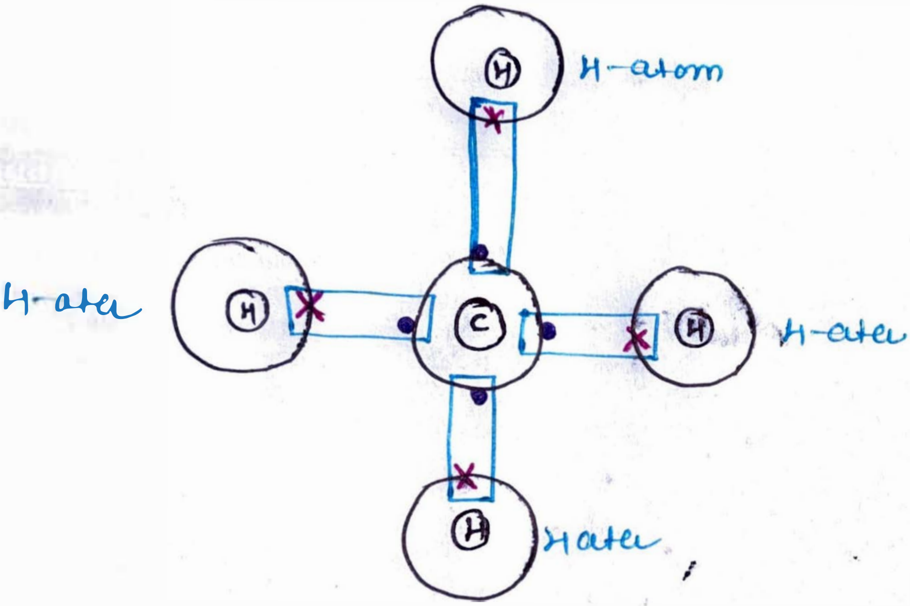
## 4→ Methane molecule (Non-Polar Compound)→

atom Involve- C, H

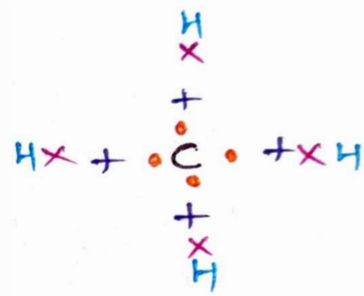
Electronic configuration- $_6C$ [2,4], $_1H$ [1]

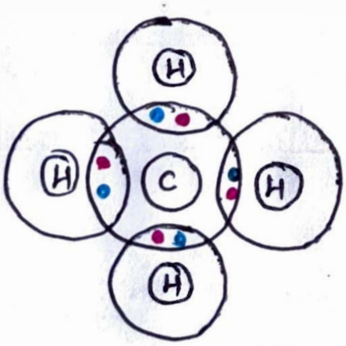
Nearest noble gas- Ne (2,8), He [2]

to attain the stable electronic configuration of the Nearest noble gas, C needs 4 $e^-$s and H needs 1 $e^-$

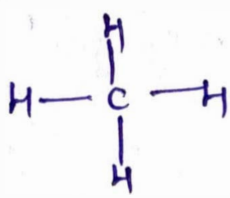



*e- Dot Structure*




Methane molecule




4-single Covalent Bonds

# A $CH_4$ molecule contains a total 5 atoms, i.e. one atom of C & 4 atoms of H.