
Kon kehta hai
11vi ki marks
matter nhi krte ?

LIVE

# CLASS 11th CHEMISTRY

## BACKLOG MARATHON

SERIES

1. Some Basic Concepts of Chemistry
2. Structure of an Atom
3. Classification of Elements & Periodicity of Properties
4. Chemical Bonding & Molecular Structure



BACKLOG MARATHON SERIES

# Class 11th | Chemistry

## Some Basic
## Concepts of chemistry

## Chapter 1

7 Marks

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# SCIENCE

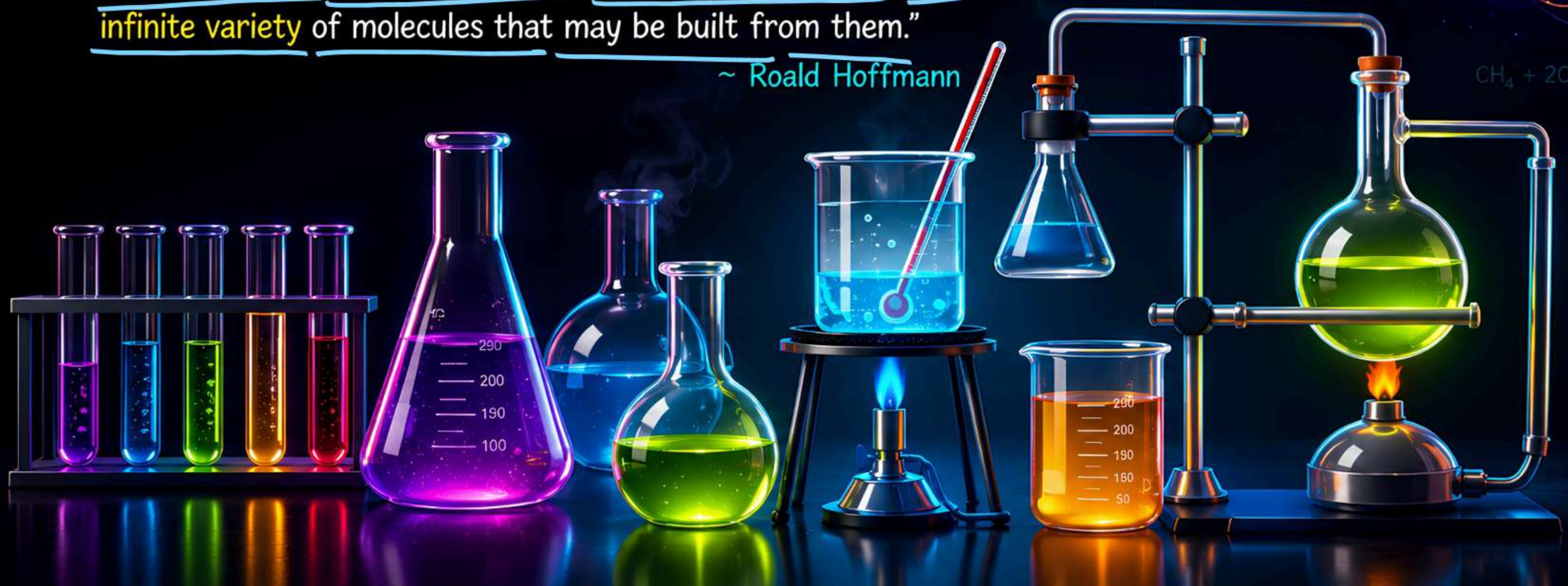
Science is the study of the natural and physical world through **observation, experimentation, and analysis**. It's a systematic way of building knowledge about the world around us.

**Latin word - Scientia** = knowledge

# DEFINITION OF CHEMISTRY

"Chemistry is the science of molecules & their transformations. It is the science not so much of the hundred elements but of the infinite variety of molecules that may be built from them."

~ Roald Hoffmann



# Nature of Matter

Temperature
Pressure.

Matter is defined as anything which occupies space and has mass, e.g., wood, book, water, air, petrol, plastics, iron etc. are all composed of matter.



# Classification of Matter

## 1. Physical classification

- In physical classification, matter is classified as solid, liquid & gas.
- The three states of matter are interconvertable. It may be done by changing the temperature and pressure.



## 2. Chemical classification

- In chemical classification, matter is classified as pure substances and mixtures.
- Materials found in nature are either single substances or consists of two or more substances.
- A sample containing only one substance is called pure substance.
- Samples containing more than one substance are not pure. They are called mixtures.

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# Matter

- Mixtures
  - Homogeneous mixtures
  - Heterogeneous mixtures
- Pure Substances
  - Elements
    - Metals
    - Non Metals
    - Metalloids
  - Compounds

$H_2O$ 1 : 8

# Measurement of Physical Properties

The International System of Units SI system is a modification of metric system and has seven base units pertaining to the seven fundamental scientific quantities.

| Property | Unit | Abbreviation |
|---|---|---|
| Length | meter | m |
| Mass | kilogram | kg |
| Time | seconds | S |
| Amount | mole | mol |
| Temperature | kelvin | K |
| Electric current | ampere | A |
| Luminous intensity | candella | cd |

Killian

Fundamental

mil K bnegi

Derived units

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# Fundamental & Derived Units

Derived

Fundamental units are independent units of measurement, while derived units are combinations of these fundamental units.

| S.no | Physical quantity | Expression | Unit |
|---|---|---|---|
| 1 | Area | length x breadth | $m^2$ |
| 2 | Volume | area x height | $m^3$ |
| 3 | Ddensity | mass / volume | $Kgm^{-3}$ |
| 4 | Velocity | displacement / time | $ms^{-1}$ |
| 5 | Momentum | mass / velocity | $Kgms^{-1}$ |
| 6 | Acceleration | velocity /time | $ms^{-2}$ |
| 7 | Force | mass/acceleration | $kgms^{-2}$ or N |
| 8 | Pressure | force/area | $Nm^{-2}$ or Pa |
| 9 | Energy/(work) | force x distance | Nm or J |
| 10 | Surface Tension | force/length | $Nm^{-1}$ |

# Prefix in SI Units

| FACTORS | PREFIX | SYMBOL | DECIMAL |
|---|---|---|---|
| $10^{24}$ | yotta | Y | 1000 000 000 000 000 000 000 000 |
| $10^{21}$ | zetta | Z | 1000 000 000 000 000 000 000 |
| $10^{18}$ | exa | E | 1000 000 000 000 000 000 |
| $10^{15}$ | peta | P | 1000 000 000 000 000 |
| $10^{12}$ | tera | T | 1000 000 000 000 |
| $10^{9}$ | giga | G | 1000 000 000 |
| $10^{6}$ | mega | M | 1000 000 |
| $10^{3}$ | kilo | k | 1000 |
| $10^{2}$ | hecto | h | 100 |
| $10^{1}$ | deca | da | 10 |

| FACTORS | PREFIX | SYMBOL | DECIMAL |
|---|---|---|---|
| $10^{-1}$ | deci | d | 0.1 |
| $10^{-2}$ | centi | c | 0.01 |
| $10^{-3}$ | milli- | m | 0.001 |
| $10^{-6}$ | micro | μ | 0.000 001 |
| $10^{-9}$ | nano | n | 0.000 000 001 |
| $10^{-12}$ | pico | p | 0.000 000 000 001 |
| $10^{-15}$ | femto | f | 0.000 000 000 000 001 |
| $10^{-18}$ | atto | a | 0.000 000 000 000 000 001 |
| $10^{-21}$ | zepto | z | 0.000 000 000 000 000 000 001 |
| $10^{-24}$ | yocto | y | 0.000 000 000 000 000 000 000 001 |

# Measurement

**Mass & Weight** — Mass of a substance is the amount of matter present in it, while weight is the force exerted by gravity on an object.

Same.

kg, g

**Volume** — Volume is the amount of space that a substance occupies. Volume of a liquid or solution can measured by graduated cylinder, burette, pipette etc.

L, mL

**Density** — Density of the substance is its amount of mass per unit volume.

$$\rho = \frac{m}{V}$$

**Temperature** — Measure of degree of hottness & coldness

Units - °C in F & K

$$0^\circ C = 273.15 K$$

$$^\circ F = \frac{9}{5}(^\circ C) + 32$$

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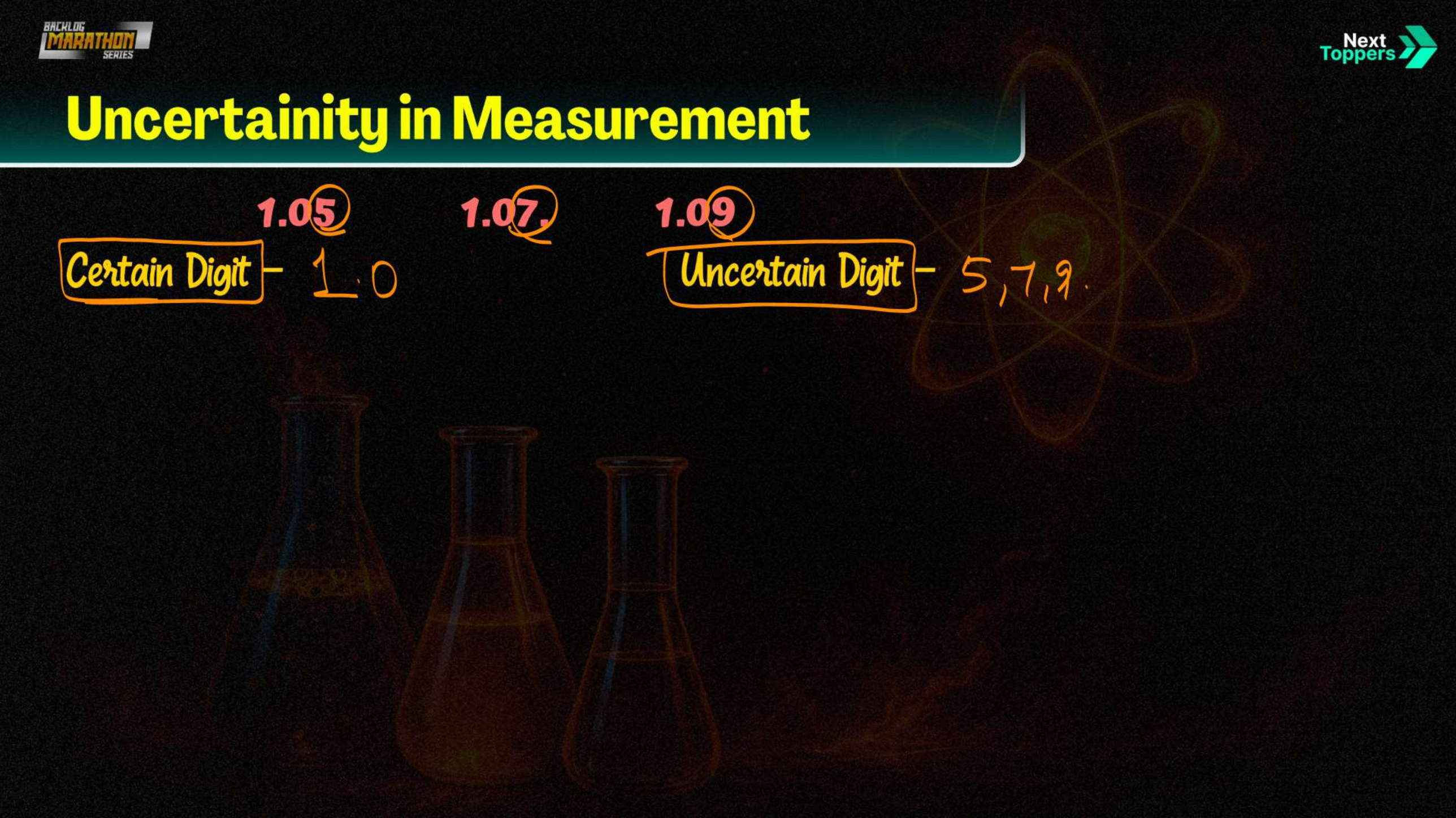
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# Uncertainity in Measurement

1.05 1.07 1.09

Certain Digit - 1.0

Uncertain Digit - 5, 7, 9.



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# Uncertainity in Measurement

## Accuracy

Accuracy describes the closeness of a single measurement to its true value. True value k aas pas value ana




## Precision

Precision describes how closely the different measurements of a given quantity agree with one another

Apas meinsame value ana




True Value

5 cm pen

Length?

Accuracy

Herman = 4.8 cm.

Ayushi = 4.99 cm

Khushboo = 4.7 cm

Values Apas mein closeness.

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# Precision & Accuracy

Lets assume true value ofsome output is 3.

| Student | Round1 output | Round2 output | Mean | Is Precise | Is Accurate |
|---|---|---|---|---|---|
| Hema | 2.80 | 3 | 2.90 | No | No |
| Rekha | 2.9 | 2.90 | 2.905 | Yes | No |
| Jaya | 3.01 | 2.99 | 3 | Yes | Yes |
| Sushma | 2.90 | 3.10 | 3 | No | Yes |

# Significant Figures

## Definition

The significant figure in a number is the sum of certain digits and one uncertain (doubtful) digit.




More Significant figures, more will be the accuracy of answers.

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# Rules to Determine Significant Figures

1. All non-zero digits are significant. 2356 → 4 S.F. ; 9237.692 → 7 S.F.
2. All the zeros between two non-zeros digits are significant figures. 2.304 → 4 S.F
3. Ending zeros in the right of the decimal are significant figures. 2.390 → 4 S.F.
4. Initial zeros are insignificant figures. 0.324 → 3 S.F ; 003.2 → 2 S.F.
5. Exact number have infinite number of significant figures. 20 chairs → ∞ S.F.
6. Power of 10 is insignificant. $2.3 \times 10^{4}$ → 2 S.F.

9.02437 → 6 S.F.

93.3000 → 6 S.F.

**Multiplication / Division** → Ques mein jis mein sbse km S.F. honge → Ans mein bhe utne he S.F honge.

2.1 × 3.32 → 6.972 → 6.9 → 2 S.F.

2 S.F. (2.1) 3 S.f (3.32)

**Addition Subtraction** → Ques mein jis mein decimal k bd sbse km s.f honge
Ans mein bhe decimal k bd utne he S.F. honge.

2.345 + 1.72 → 4.065 → 4.06

3 S.f (2.345) 2 S.f. (1.72)

**Roundoff:-** upto 2 S.F.

5.372 → 5.4

5.322 → 5.3

odd 5.35 → 5.4 (+1)

even 5.25 → 5.2.

# Laws Of Chemical Combinations

The chemical reactions take place according to the certain laws.
These are called laws of chemical combination.

1. Law of conservation of mass
2. Law of definite or constant proportions
3. Law of multiple proportions
4. Gay Lussac's Law
5. Avogadro's Law

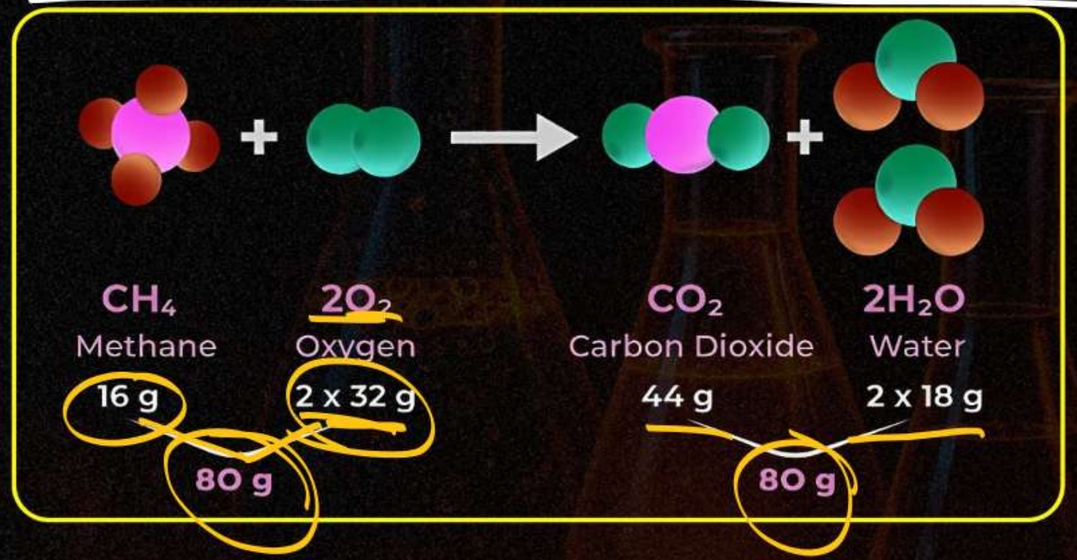
# 1. Law of Conservation of Mass

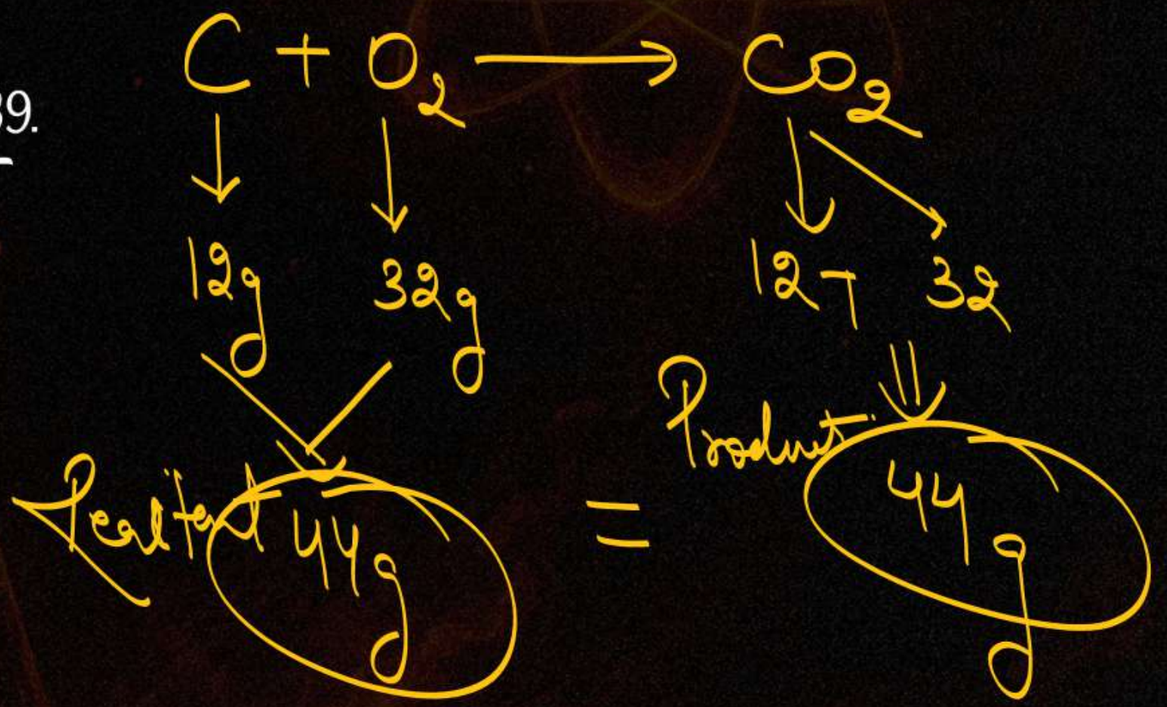
In a physical or chemical change, the total mass of the products is equal to the total mass of the reactants.

**OR**

Matter can neither be created nor destroyed.

This Law was put forth by Antoine Lavoisier in 1789.



# 2. Law of Definite or Constant Proportions

## Definition

$CuCO_3$ ≠

A chemical compound always contains same elements and their proportion by mass is always definite and constant. This law was given by Proust.





$H_2O$

H : O

1 : 8 → fixed.

Khushnuma – R.O. pani

Saumya – Sea water

Pranav – Tanki

# Prarambh 2.0 → NEXT 500

Prarambh
↓
Students

Prarambh Plus → NEXT 1500

Description → Enroll

# 3. Law of Multiple Proportions

## Definition

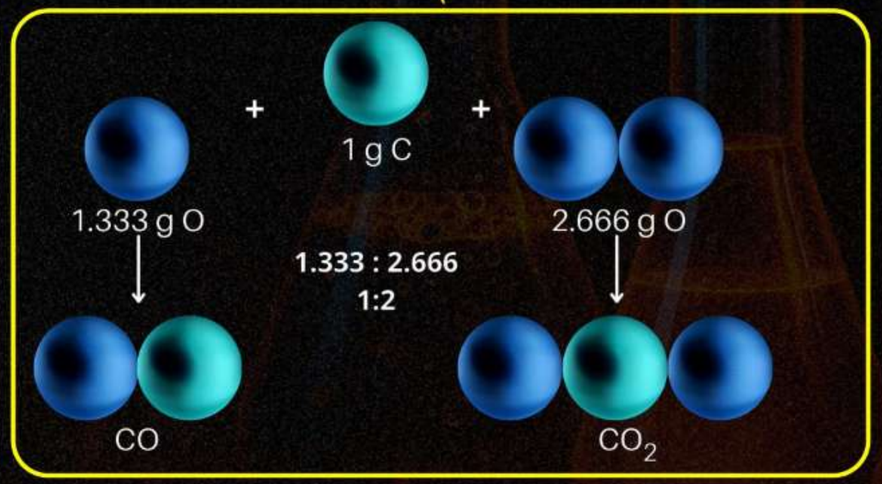
$N + O_2 \rightarrow NO_2$

$N + O_2 \rightarrow N_2O_3$

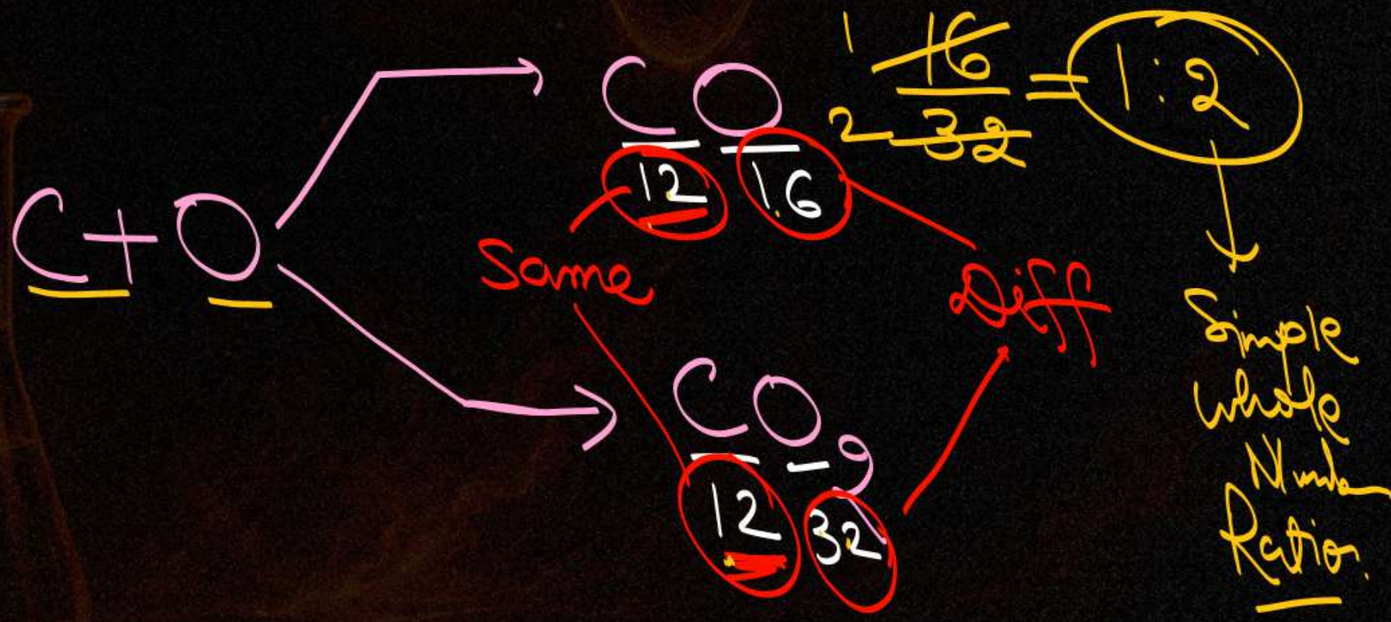
$N + O_2 \rightarrow N_2O_5$

When two elements combine to form two or more compounds, then the different masses of one element which combine with the same mass of other element are in simple ratio of whole numbers.

This law was given by Dalton in 1803.




∴



# 4. Gay-Lussac Law of Gaseous Volumes

## Definition

When gases react they do so in simple multiple ratio of volumes and if the products formed are also gases, all the volumes bear a simple ratio provided all the volumes are measured at same temperature and pressure.

This law is applicable to gaseous reactions.

It was given by Gay Lussac in 1808 .

$$\underset{\text{2 volume}}{2H_2\,(g)} + \underset{\text{1 volume}}{O_2\,(g)} \longrightarrow \underset{\text{2 volume}}{2\,H_2O\,(g)}$$

2 : 1 ⇒ S.W.R.

2 : 1 : 2 ⇒ S.W.R.

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# 5. Avogadro's Law




## First definition:

Equal volume of all gases under the same conditions of temperature and pressure contain equal number of molecules.

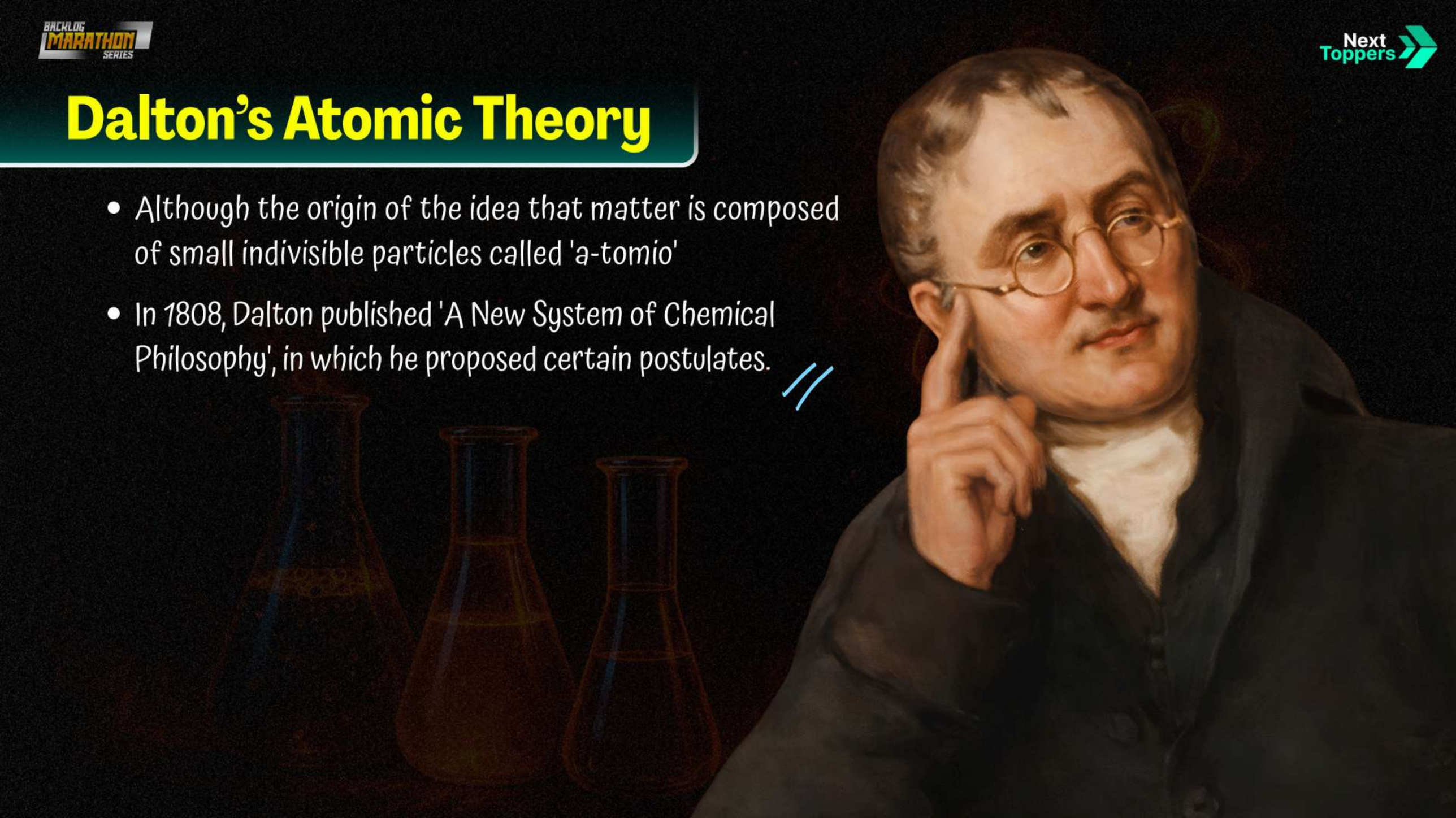
## Second definition:

The volume of a gas at constant temperature and pressure is proportional to the number of moles or molecules of the gas present.

$$\underset{\substack{2 \text{ mole} \\ 2\times6.02\times10^{23} \\ \text{molecules}}}{2H_2\,(g)} + \underset{\substack{1 \text{ mole} \\ 1\times6.02\times10^{23} \\ \text{molecules}}}{O_2} \longrightarrow \underset{\substack{2 \text{ mole} \\ 2\times6.02\times10^{23} \\ \text{molecules}}}{2\,H_2O\,(g)}$$

# Dalton's Atomic Theory

- Although the origin of the idea that matter is composed of small indivisible particles called 'a-tomio'
- In 1808, Dalton published 'A New System of Chemical Philosophy', in which he proposed certain postulates.



# Postulates of Dalton's Atomic Theory

1. Matter is made up of extremely small indivisible and indestructible ultimate particles called atoms.
2. Atoms of the same element are identical in all respects i.e. in shape, size, mass and chemical properties.
3. Atoms of different elements are different in all respects and have different masses and chemical properties.
4. Atom is the smallest unit that takes part in chemical combinations.
5. Atoms of two or more elements combine in a fixed ratio to form compound.
6. Atom can neither be created nor be destroyed during any physical or chemical change.

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# Atomic & Molecular Mass

## Atomic mass

$Z <$ even $\rightarrow 2 \times Z$, odd $\rightarrow 2 \times Z + 1$.

$Na \rightarrow 11 \times 2 + 1 = 23$

$Al \rightarrow 13 \times 2 + 1 = 27$

It is the total mass of an atom, which is the sum of the masses of its protons & neutrons.

1 a.m.u = 1 g

## Molecular Mass

$H_2SO_4 \rightarrow 1 \times 2 + 32 + 16 \times 4 \Rightarrow 2 + 32 + 64 = 98\ u$

The sum of the atomic masses of all the atoms in a molecule of a substance is the molecular mass.

## Formula Mass

$NaCl \rightarrow 23 + 35.5 \rightarrow 58.5\ u$.

The formula mass of a substance is defined as the sum of the atomic masses of constituent atoms in an ionic compound. This is generally used for ionic compounds which do not contain discrete molecules, but ions as their constituent units.

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# Atomic & Molecular Mass

## Average atomic Mass

The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance

| Isotopes | Relative abundance | AMU |
|---|---|---|
| C 12 | 98.892 | 12 |
| C 13 | 1.108 | 13.0035 |
| C 14 | $2*10^{-10}$ | 14.0037 |

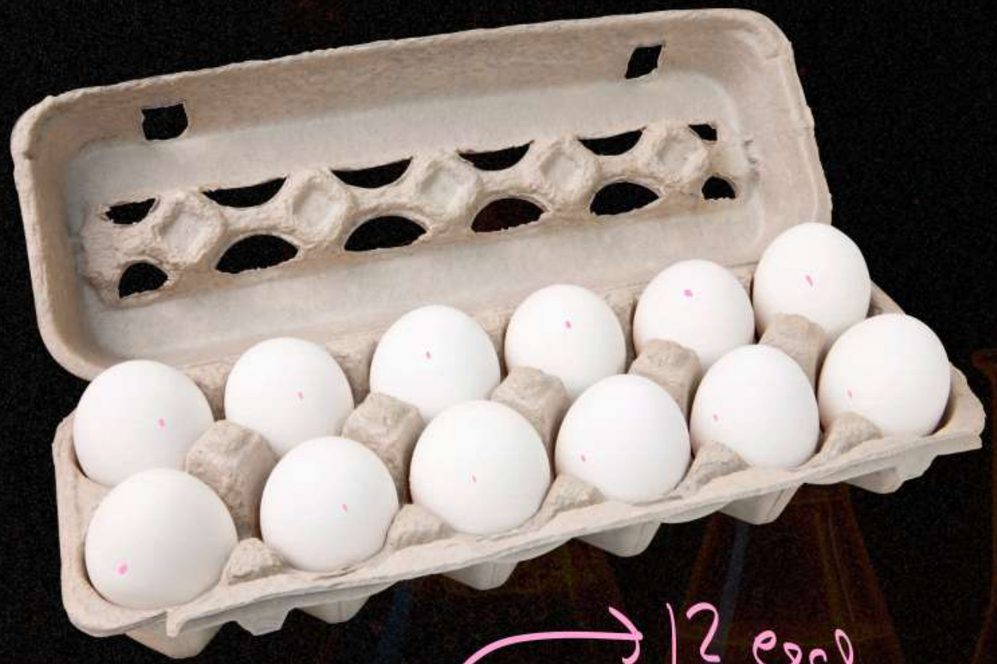
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# Mole Concept

→ Number

1 mole ⇒ $6.023 \times 10^{23}$ → $N_A$ =



→ 12 eggs.

I dozen eggs



→ 2 shoes

I pair of shoes

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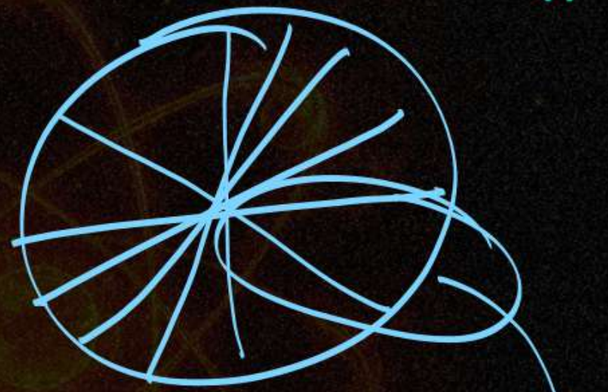
# Formulas

Mole = $n = \frac{w}{M}$

$n = \frac{\text{given particle}}{N_A}$

$n = \frac{V}{22.7}$

Questions



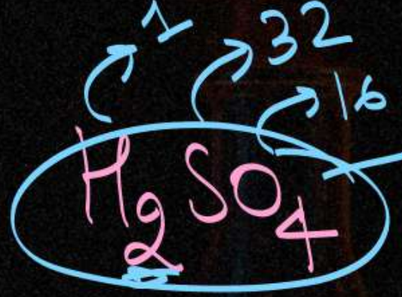
$1\,u \leftarrow \frac{1}{12^{th}}$ C-atom

# Percentage Composition

Percentage composition is the percentage of each element in a compound by mass:-

$$\%composition = \frac{n \times M(element)}{M(element)}$$

$M(element)$ → Total Mass




$2+32+64 = 98$

$H\% = \frac{2 \times 1}{98} \times 100$

$\%\ S = ?$

$\%\ S = \frac{32 \times 100}{98}$

$O\% \rightarrow \frac{4 \times 16}{98} \times 100$

# Emperical Formula

## Empirical formula

is the formula of a compound which shows the - simplest whole number ratio between the atoms of the elements in the compou

| Element | Percentage | Atomic mass | Moles of atom | Atomic ratio | Simplest ratio |
|---|---|---|---|---|---|
| Na | 14.31 | 23 | $\frac{14.31}{23}=0.622$ | $\frac{0.622}{0.311}=2$ | 2 |
| S | 9.97 | 32 | $\frac{9.97}{32}=0.311$ | $\frac{0.311}{0.311}=1$ | 1 |
| H | 6.22 | 1 | $\frac{6.22}{1}=6.22$ | $\frac{6.22}{0.311}=20$ | 20 |
| O | 69.5 | 16 | $\frac{69.5}{16}=4.34$ | $\frac{4.34}{0.311}=13.95$ | 14 |

Let

| 1 | 2 | 3 | 4 | 5 |
|---|---|---|---|---|
| Element | Percentage comp. | Mass | Moles (Box 4 mein ×2 ÷ 1 Box) | S.R. (Box 4 mein se smallest no. uthao or usse box K sari values se divide karo.) |
| H | 10% | 1 | 10 ÷ 1 = 2 | 2 ÷ 2 ⇒ 1 |
| S | 60% | 32 | 60 ÷ 32 = 8 | 8 ÷ 2 ⇒ 4 |
| O | 30% | 16 | 30 ÷ 16 = 10 | 10 ÷ 2 ⇒ 5 |

E.f = $HS_4O_5$

Possible nhi hai sirf example hai

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# Emperical Formula

$M.M = 2 \times V.D.$

## Molecular formula

is the chemical formula which represents - the actual number of atoms of each element present in a molecule of the compound.

$$n = \frac{\text{Molecular weight}}{\text{Empirical formula weights}}$$

$M.F. = n \times E.F.$ (n: common)

$\text{Molecular Mass} = n \times \text{Emp. Mass}$

M.F $C_6H_{12}O_6$ → 12 : 6 : 6 → 2 : 1 : 1

E.F → Simplest Ratio → $CH_2O$ × 6

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# Stoichiometry

1) Balance
2) Poocha hai → Relation. moles
   Diya hai
3) Equate...

- Stoichiometry deals with calculations of masses and sometimes also volumes of the reactants and products involved in a chemical reaction.
- We can also say that it is nothing but basically the relationship between the amounts of reactants and the products in a chemical reaction.

2 mole

$2H_2 + 1O_2 \rightarrow 2H_2O$

| 1 MOLE HYDROGEN | 1 MOLE OXYGEN | 1 MOLE OF WATER |
|---|---|---|
| 2G OF HYDROGEN | 32G OF OXYGEN | 18G OF WATER |
| 2 ATOMS OF HYDROGEN | 2 ATOMS OF OXYGEN | 2 ATOMS OF HYDROGEN – 1 ATOMS F OXYGEN |

$2H_2 \rightarrow 1O_2$

$1H_2 \rightarrow \frac{1}{2}O_2$

# Concentration Terms

$$m\% = \frac{m_{solute}}{m_{soln.}} \times 100$$

- **Mass Percentage -** It is equal to the mass of a solute in grams present in 100 g of a solution.
- **Molarity -** It is equal to the number of gram moles of solute present in 1 litre of solution.
- **Molality -** It is equal to the number of moles of the solute present in 1 kg (1000 g) of the solvent.
- **Mole Fraction -** It is the ratio of moles of one component to the total number of moles (solvent and solute) present in the solution . it is important to know that sum of the mole fraction of all the components is always unity.

$$M = \frac{n_{solute}}{V_{soln.}(in\ L)}$$

$n = \frac{w}{M}$

unit → M → mol/L

unit : m = mol/kg

$$m = \frac{n_{solute}}{Mass_{solvent}(in\ kg)}$$

$$\chi_A = \frac{n_A}{n_A + n_B}$$

$$\chi_B = \frac{n_B}{n_A + n_B}$$

No unit





1 ROOM = 1 FAMILY

20 ROOMS 21 FAMILIES ???

20 ROOMS- LIMITING REAGENT 21 FAMILIES- EXCESS REAGENT

# Limiting Reagent

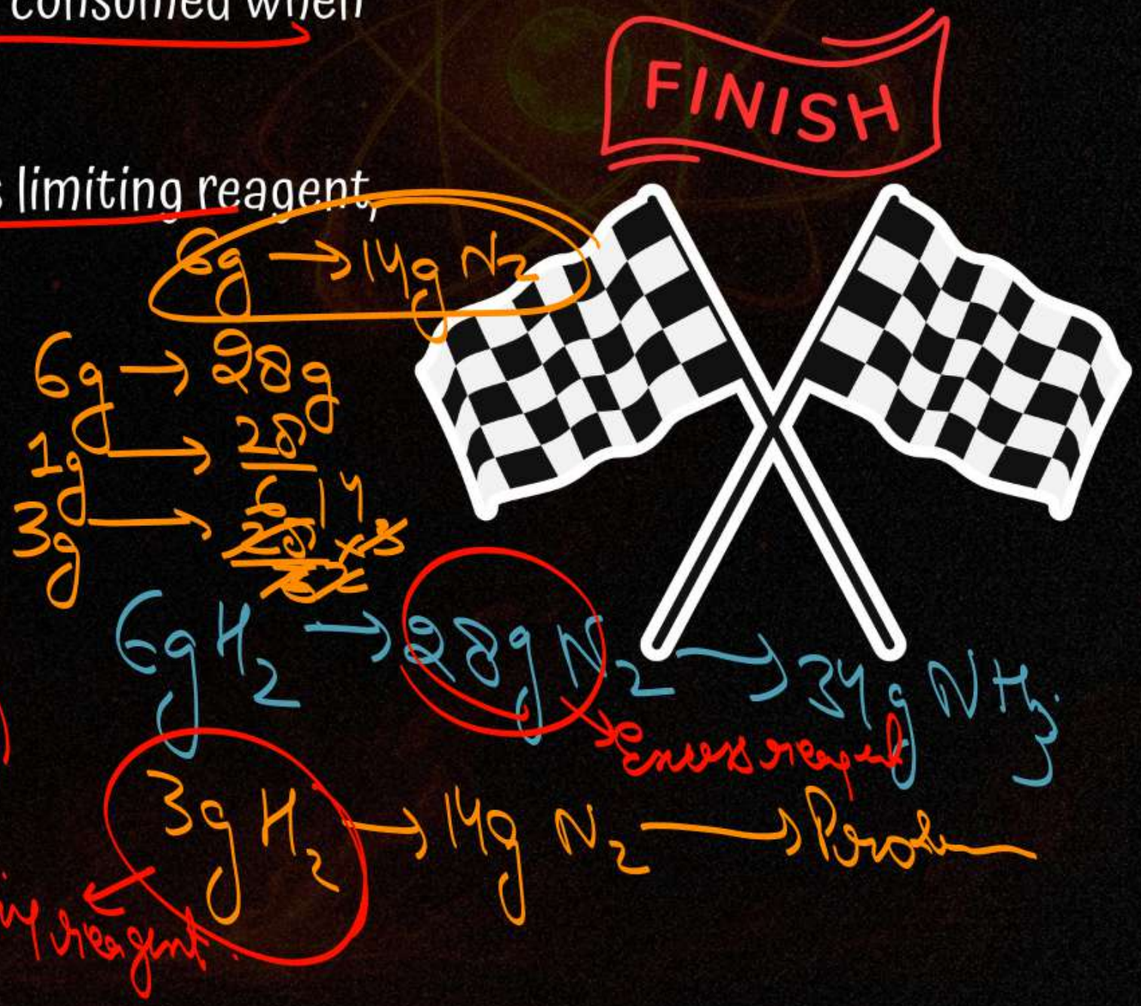
- Limiting reagent is the substance that is totally consumed when the chemical reaction is complete.
- The amount of product formed is limited by this limiting reagent, so the reaction cannot continue without it.

Absence of limiting reagent

$$\underset{6\text{ g}}{3\ H_2} + \underset{28\text{ g}}{N_2} \longrightarrow \underset{34\text{ g}}{2\ NH_3}$$

Presence of limiting reagent

$$\underset{\substack{3\text{ g}\\\text{Limiting reagent}}}{1.5\ H_2} + \underset{28\text{ g}}{N_2} \longrightarrow \underset{17\text{ g}}{NH_3} + \underset{\substack{14\text{ g remaining}\\\text{Excess reagent}}}{N_2}$$



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# Difference Between Excess Reagent & Limiting Reagent

| Limiting reagent | Excess reagent |
|---|---|
| Limiting reagent is the reactant of a particular chemical reaction that limits the formation of the product | Excess reagent is the reactant that is present in excess in a reaction mixture |
| Completely consumed during a reaction | Not completely consumed during a reaction |
| Not present at the end of the reaction | Some amount is present at the end of the reaction |
| Limits the amount of product that is formed from a reaction | Has no effect on the product formed from the chemical reaction |


Apne
itne acche se
pdha
isliye meri taraf se
5 minute ka
break lelo
7 min Break
19:10
Keep Going!
Proud of you
You Did Great!
Break Time
Hard Work Today
Better Tomorrow
Future is Bright
Notes

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# Class 11th | Chemistry

## Structure of
## An Atom

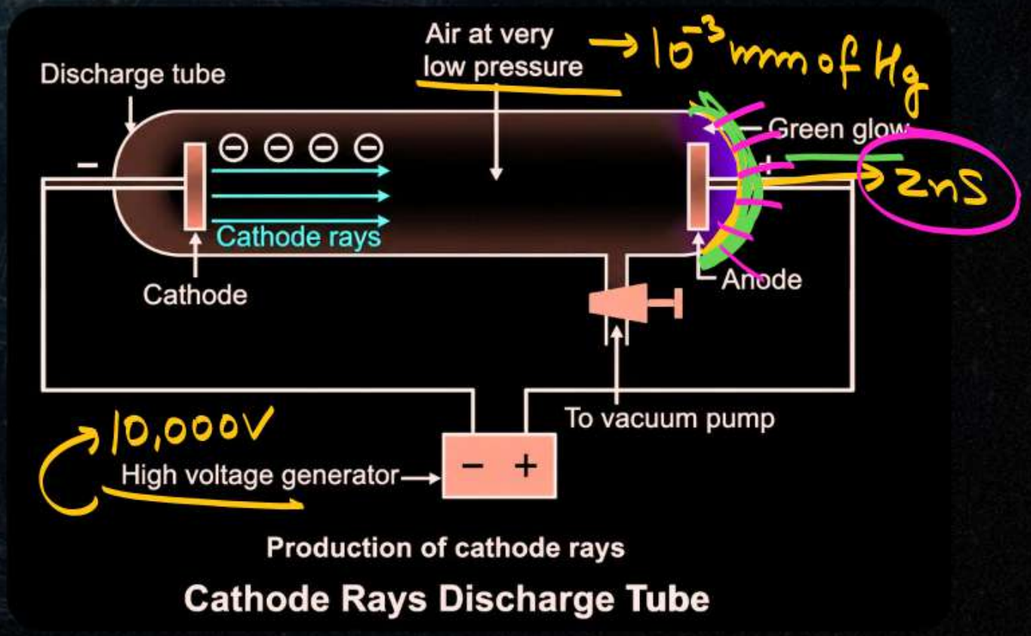
## Chapter 2

9 Marks

# Cathode Rays

**A cathode ray tube is made of glass containing two thin pieces of metal, called electrodes, sealed in it.**

- The electrical discharge through the gases could be observed only at very low pressures and at very high voltages.
- The pressure of different gases could be adjusted by evacuation of the glass tubes.



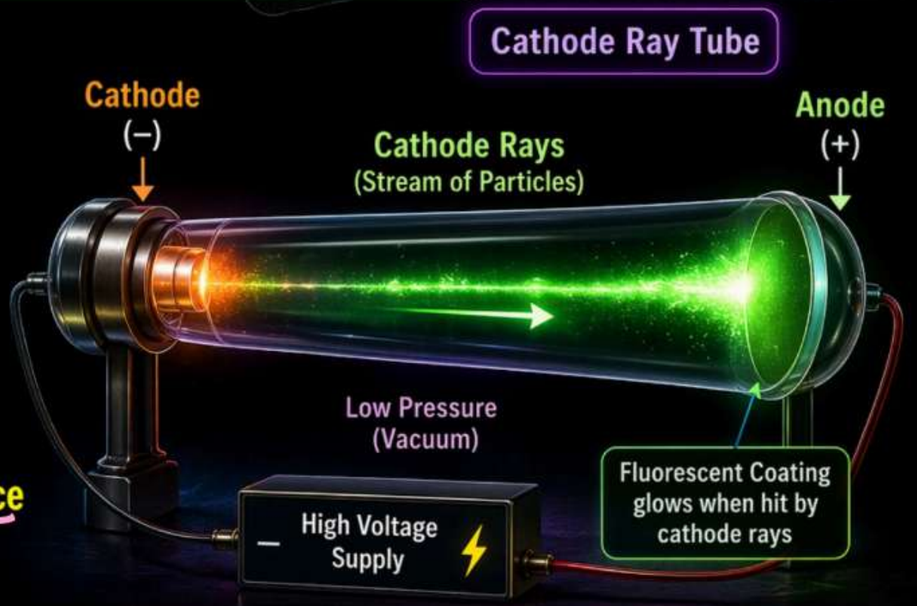

Cathode Rays Discharge Tube

# Cathode Rays

- When sufficiently high voltage is applied across the electrodes, current starts flowing through a stream of particles moving in the tube from the negative electrode (cathode) to the positive electrode (anode).
- These were called cathode rays or cathode ray particles.
- The flow of current from cathode to anode was further checked by making a hole in the anode and coating the tube behind anode with phosphorescent material zinc sulphide.
- When these rays, after passing through anode, strike the zinc sulphide coating, a bright spot is developed on the coating.

# Properties of Cathode Rays

- The cathode rays start from **cathode** and move towards the **anode**.
- These rays themselves are **not visible** but their behaviour can be observed with the help of certain kind of materials (**fluorescent** or **phosphorescent**) which glow when hit by them.
- Television picture tubes are cathode ray tubes and television pictures result due to **fluorescence** on the television screen coated with certain **fluorescent** or **phosphorescent** materials.




**Key Properties**

- Originate from Cathode (–)
- Move towards Anode (+)
- Make certain materials glow

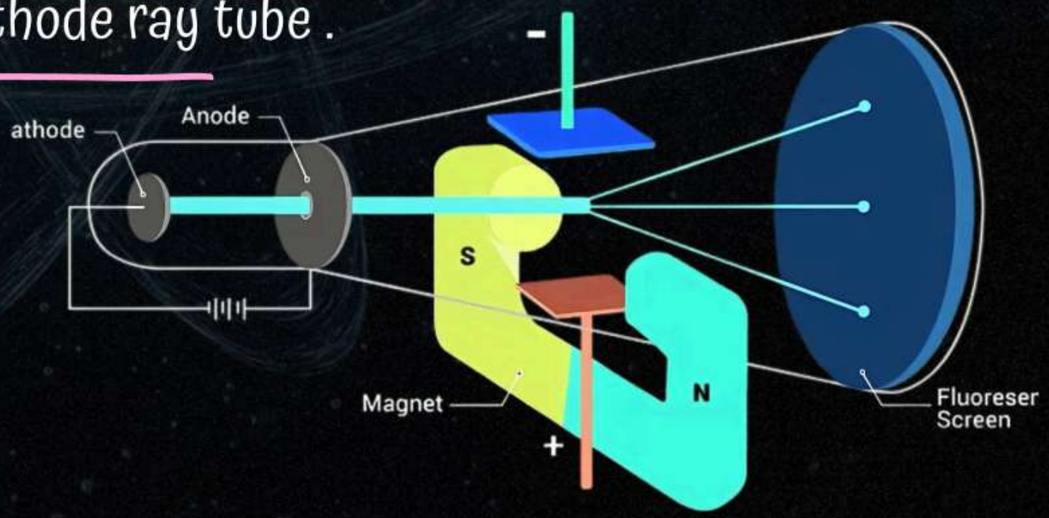
**Television Picture Tube**

An Example of Cathode Ray Tube in Daily Life



# Properties of Cathode Rays

- In the absence of electrical or magnetic field, these rays travel in straight lines .
- In the presence of electrical or magnetic field, the behaviour of cathode rays are similar to that expected from negatively charged particles, suggesting that the cathode rays consist of negatively charged particles, called electrons.
- The characteristics of cathode rays (electrons) do not depend upon the material of the electrode and the nature of the gas present in the cathode ray tube .



# Charge to Mass Ratio

$e \rightarrow -1.6 \times 10^{-19} C$

$m_e \rightarrow 9.1 \times 10^{-31} kg$ } → Same

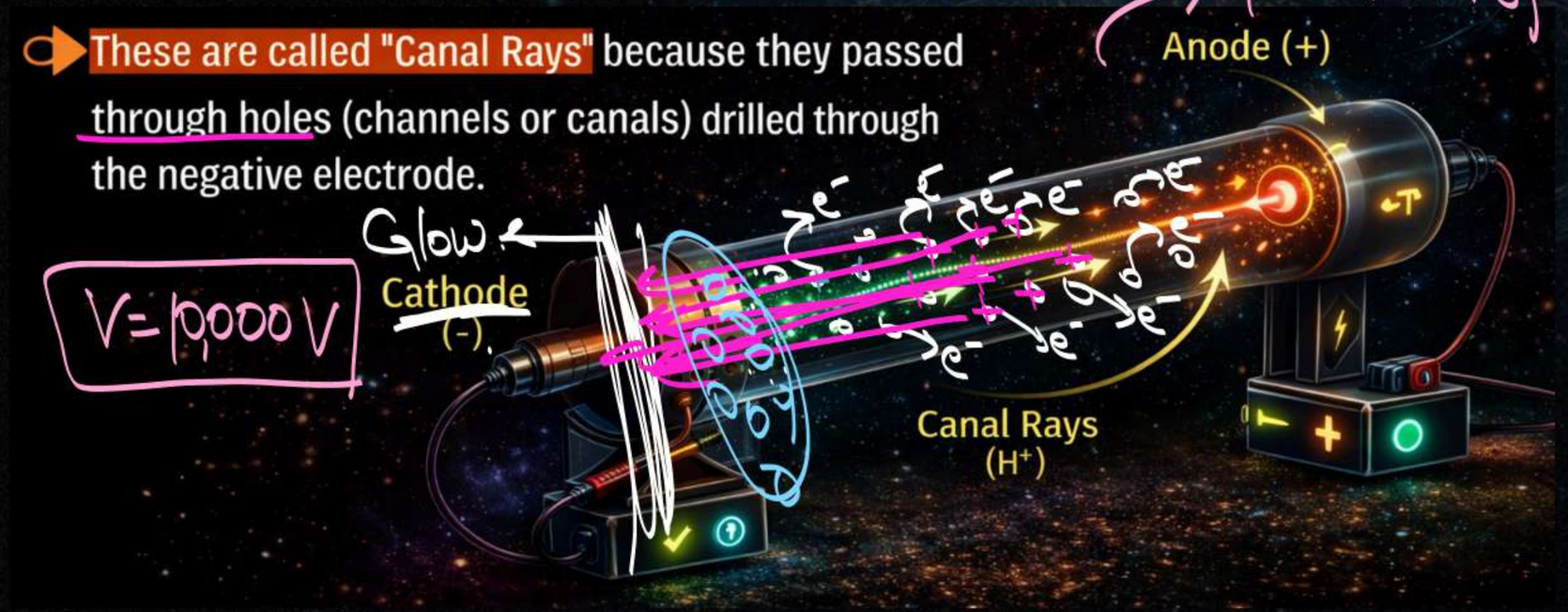
- Thomson measured charge-to-mass ratio of cathode rays by measuring their deflection under the simultaneous influence of electric and magnetic field.
- This ratio is independent of the nature of the cathode and the gas in the discharge tube.
- It proves that all gases taken were made up from negatively charged fundamental particles. Thomson called these particles corpuscles.
- Charge-to-mass ratio for anode rays depends upon the nature of gas and is lesser than e/m of electron.




$$\frac{e}{m_e} = 1.758820 \times 10^{11} Ckg^{-1}$$

# Anode Rays

- Eugene Goldstein noted streams of positively charged particles in cathode rays in 1886.
- Particles move in opposite direction of cathode rays.

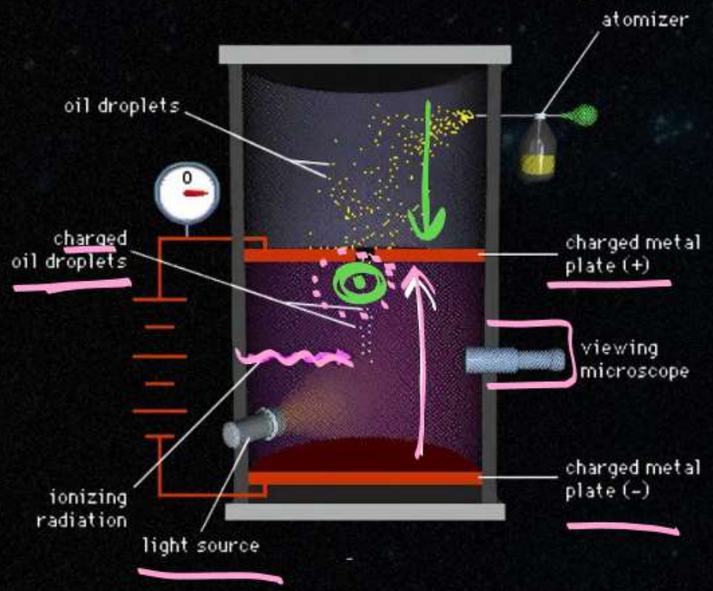


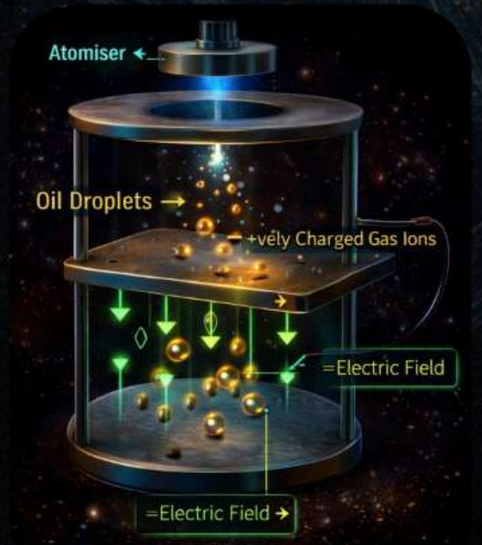
# Properties of Canal Rays

- Unlike cathode rays, mass of positively charged particles depends upon the nature of gas present in the cathode ray tube.
- These are simply the positively charged gaseous ions.
- The charge to mass ratio of the particles depends on the gas from which these originate.
- Some of the positively charged particles carry a multiple of the fundamental unit of electrical charge.
- The behaviour of these particles in the magnetic or electrical field is opposite to that observed for electron or cathode rays.

# Millikan's Oil Drop Experiment

- **The Millikan oil drop experiment, conducted by Robert Millikan in the early 1900s, was a crucial experiment that accurately determined the charge of a single electron.** By observing the motion of tiny oil droplets in an electric field, Millikan was able to calculate the elementary charge.






# Millikan's Oil Drop Experiment

## Atomiser:

Containing oil that is used to spray fine oil droplets into the top chamber of the apparatus. Some oil droplets fell inside the chamber through the hole.

- Charge on an electron
- Force of Gravity (F) or Weight = mg
- Force due to electric field (F=) = qE
- Now, oil droplet suspended in air when F = Fe

## X-Ray: ON

- Air is ionised (+vely charged gas ions) and -vely charged electrons.
- Electrons surrounds the oil droplets and making them -vely charged.
- The voltage across the plates was adjusted such that the oil droplets suspend in air.

$n \times -1.6 \times 10^{-19}$ → charge on electron

**Millikan observed that oil droplets have different charges but each was a whole number multiple of -1.6 X 10-19 C. He considered it to the fundamental charge.**

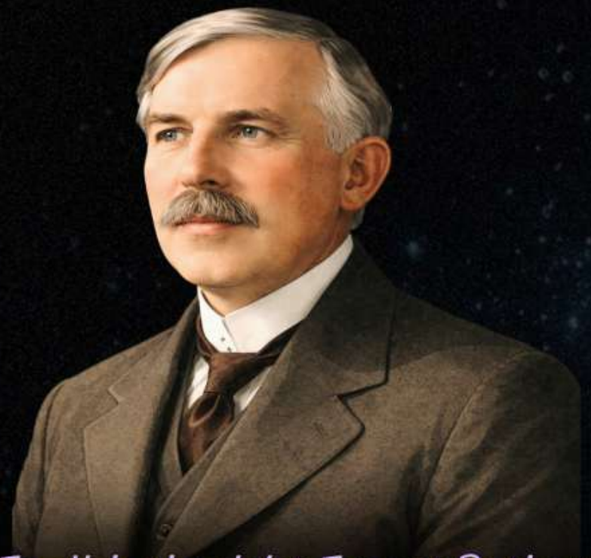
# Thompson Model of Atom



In 1897, the English scientist J.J.Thomson provided the first hint that an atom is made of even smaller particles.

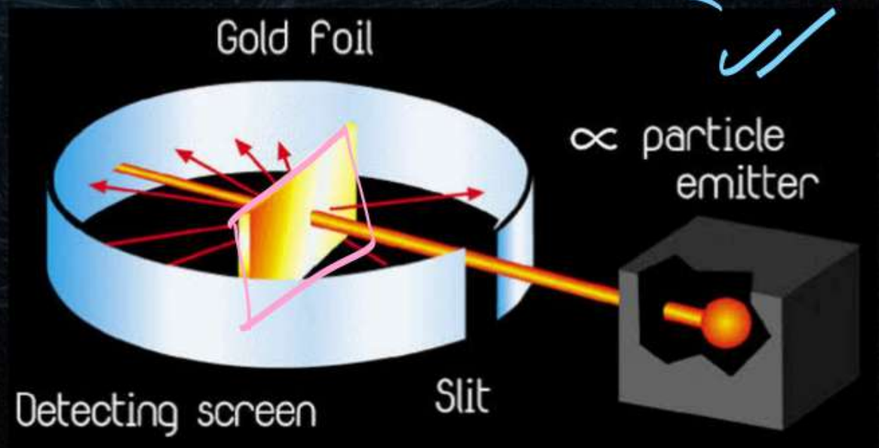
- **He proposed a model of the atom that is sometimes called the "Plum Pudding" model.**
- Atoms were made from a positively charged sphere with negatively charged electrons embedded in it, like raisins in a puddi ng.
- The negative and positive charges are equal in magnitude. So, the atom as a whole is electrically stable.

# Rutherford Model of Atom



In 1908, the English physicist Ernest Rutherford was hard at work on an experiment that seemed to have little to do with unraveling the mysteries of the atomic structure.
This is also known as Geiger-Marsden Experiment or Gold Foil Experiment

- Rutherford's experiment Involved firing a stream of tiny positively charged alpha (α) particles at a thin sheet of gold foil.
- The expected result was that the a-particles would be deflected by the subatomic particles in gold atoms. Since a - particles were much heavier than protons, he did not expect to see larger deflections.

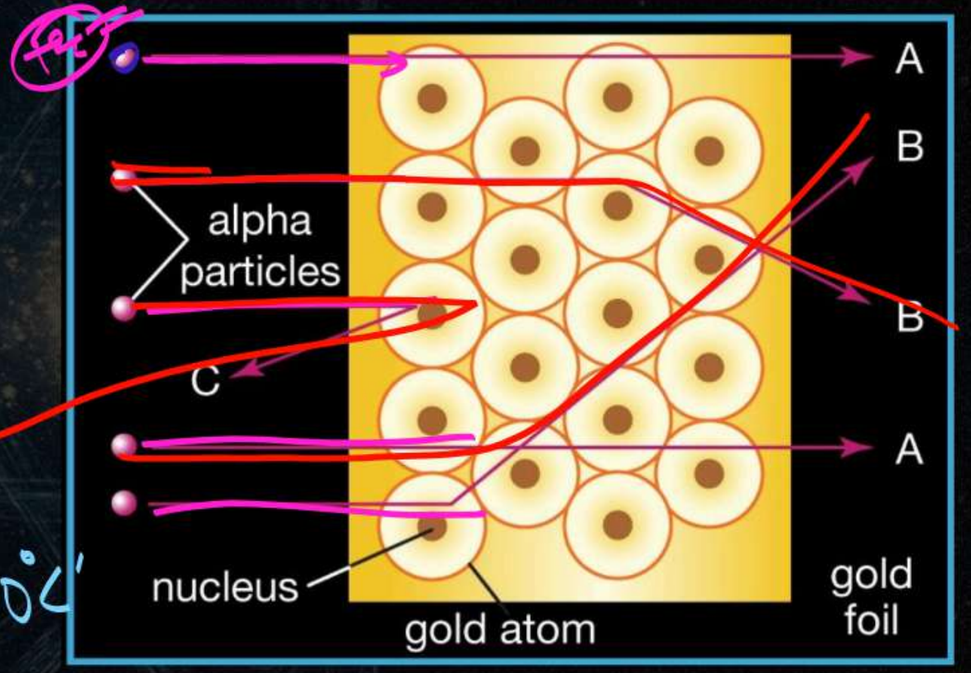


# Postulates of Rutherford Atomic Model

1) Most of the α-particle ⟶ straight
   ↳ Reason → Most Portion → Atom → Empty.

2) Some α-particles → deflected → Small ∠'
   ↳ Reason → Some +ve → present atom.

3) Veryyy few α-particles → deflected by 180° ∠'
   ↳ whole mass of Atom → Situated very small portion ⟶ Nucleus → protons + neutrons



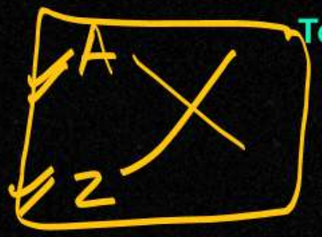
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Next Toppers

$Z = p = e$

$A = p + n$

$n = A - Z$




# Atomic Number & Mass Number

| Name of the elements | Symbol of the elements | Atomic mass |
|---|---|---|
| Hydrogen | H | 1 amu |
| Helium | He | 4 amu |
| Lithium | Li | 7 amu |
| Berylium | Be | 9 amu |
| Boron | B | 11 amu |
| Carbon | C | 12 amu |
| Nitrogen | N | 14 amu |
| Oxygen | O | 16 amu |
| Flourine | F | 19 amu |
| Neon | Ne | 20 amu |

| Name of the elements | Symbol of the elements | Atomic mass |
|---|---|---|
| Sodium | Na | 23 amu |
| Magnesium | Mg | 24 amu |
| Aluminium | Al | 27 amu |
| Silicon | Si | 28 amu |
| Phosphorous | P | 31 amu |
| Sulphur | S | 32 amu |
| Chlorine | CI | 35 amu |
| Argon | Ar | 40 amu |
| Potassium | K | 39 amu |
| Calcium | Ca | 40 amu |

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## Isotopes → Atoms of Same Element having same 'Z' → diff → 'A'

Diff A → (1) (2) (3) — $_1H$, $_1H$, $_1H$

Same Z → (1) (1) (1)

A → (12) (13) (14) — $_6C$, $_6C$, $_6C$

Z → (6) (6) (6)

## Isobars → Atoms of Different Element having Same 'A' Diff 'Z'

Same A → (40) (40) — $_{18}Ar$, $_{20}Ca$

Diff 'Z' → (18) (20)

Isotones → Same number of neutrons

## Isoelectronic → Same number of Electrons

$Na^+$ → $10e^-$ , $F^-$ → $10e^-$ , $Mg^{+2}$ → $10e^-$

# Bohr's Atomic Model

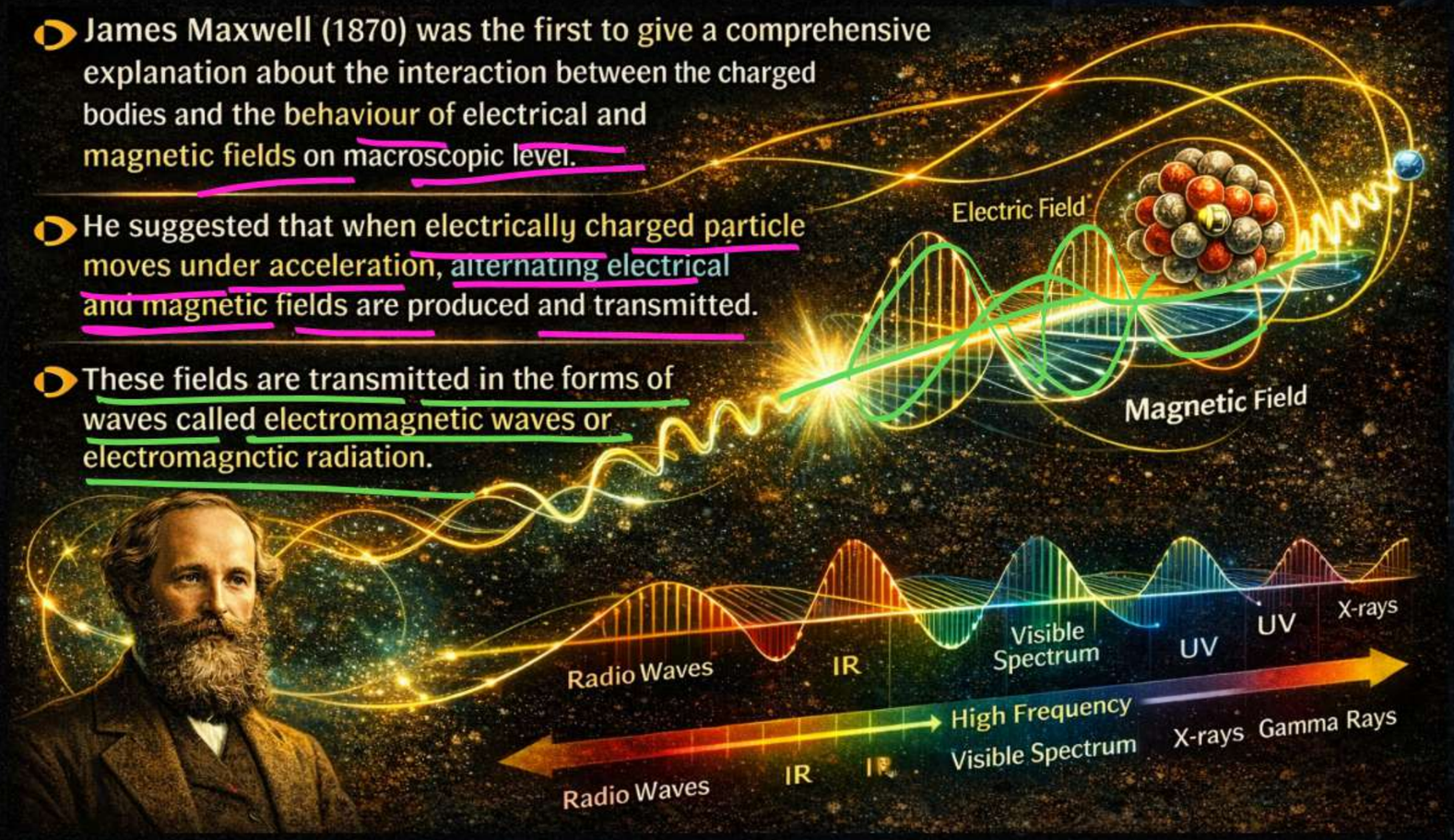


In 1913, the Danish scientist Niels Bohr proposed an improvement.
In his model, he placed each electron in a specific energy level.

- Dual character of the electromagnetic radiation which means that radiations possess both wave like and particle like properties, and Experimental results regarding atomic spectra.
- In an atom, it was realised that the attraction between the electrons and the protons should make the atom unstable.
- Bohr proposed a model in which the electrons would stably occupy fixed orbits, as long as these orbits had special discrete locations.

# Wave Nature of Electromagnetic Radiation

- James Maxwell (1870) was the first to give a comprehensive explanation about the interaction between the charged bodies and the behaviour of electrical and magnetic fields on macroscopic level.
- He suggested that when electrically charged particle moves under acceleration, alternating electrical and magnetic fields are produced and transmitted.
- These fields are transmitted in the forms of waves called electromagnetic waves or electromagnctic radiation.

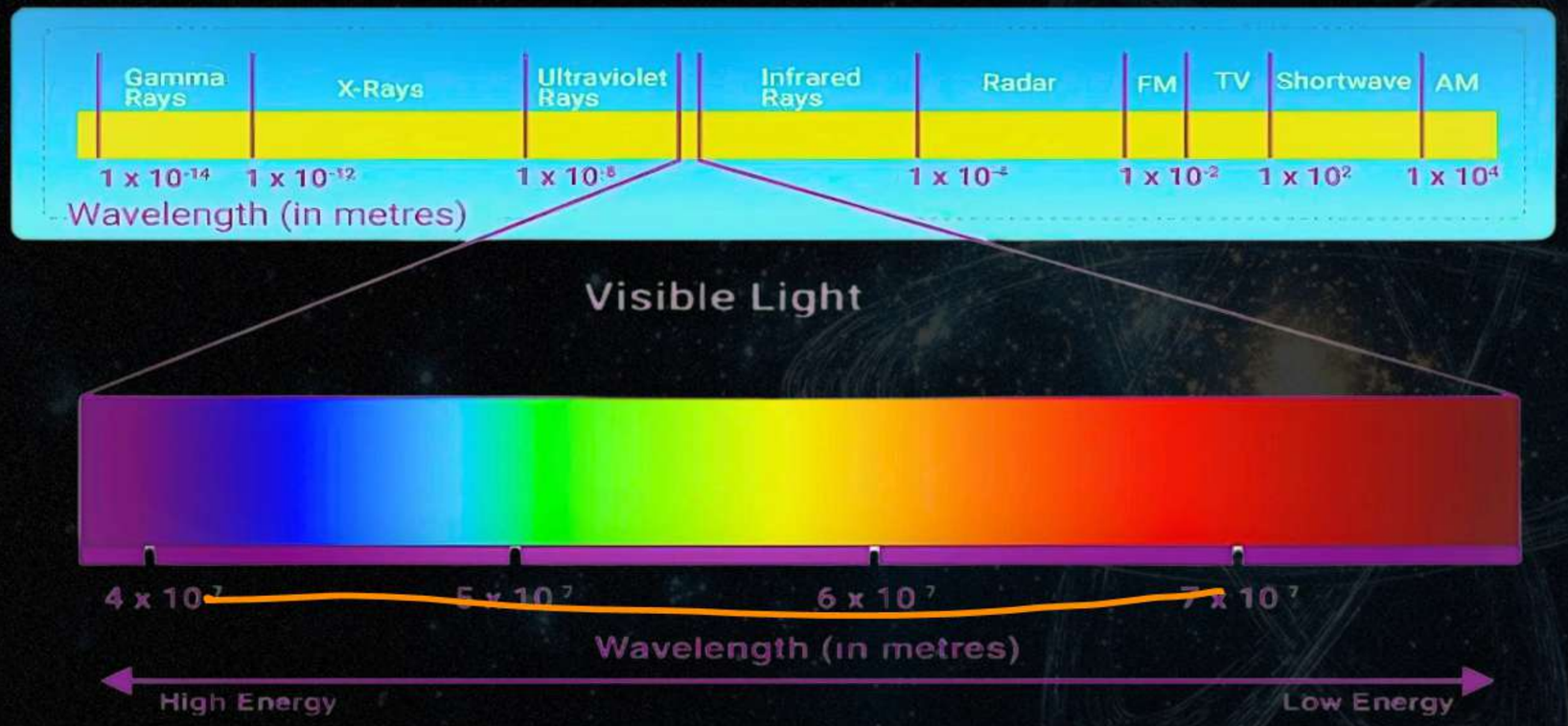






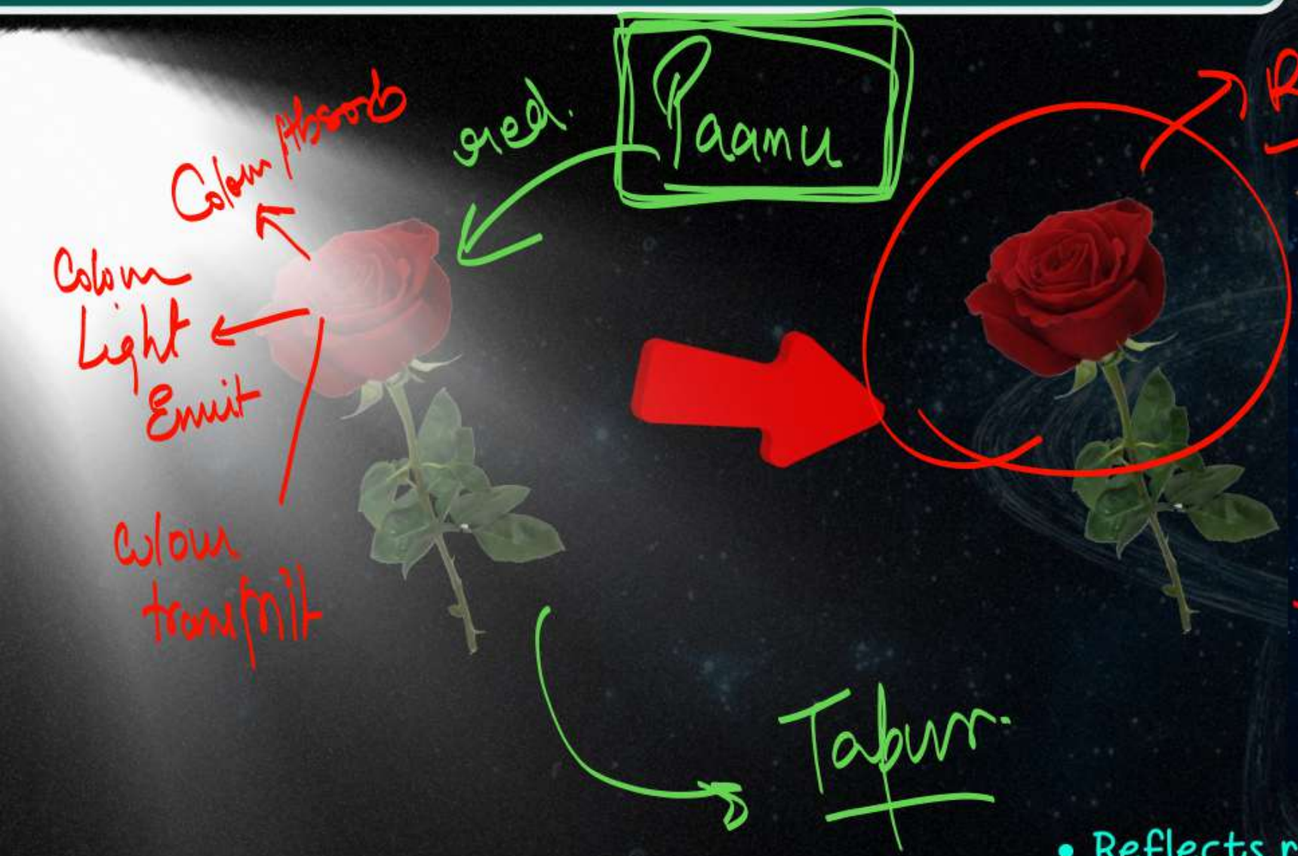
- The electric and magnetic field components of an electromagnetic wave.
- These components have the same wavelength, frequency, speed and amplitude, but they vibrate in two mutually perpendicular planes.

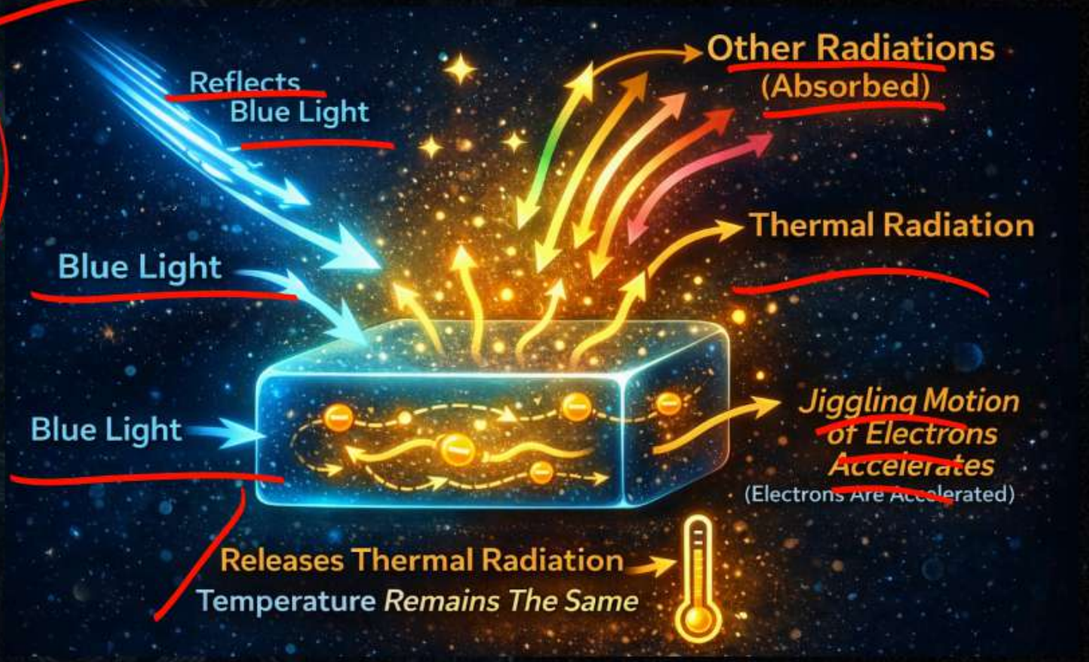
# Electromagnetic Spectrum




Wavelength decreases, Frequency increases , Penetrating power increase

# Black Body Radiation




- Reflects red light but absorbs other radiations.
- This increases jiggling motion of electron (electrons are accelerated) & it releases thermal radiation & the temperature remains the same.

# Black Body Radiation

- When an iron rod is heated in a furnace, it first turns to dull red and then progressively becomes more and more red as the temperature increases. As this is heated further, the radiation emitted becomes white and then becomes blue as the temperature becomes very high. This means that red radiation is most intense at a particular temperature and the blue radiation is more intense at another temperature. This means intensities of radiations of different wavelengths emitted by hot body depend upon its temperature.

When we heat iron all color of wavelengths are emitted but our eyes can percieve only Red colour because it has maximum wavelength.

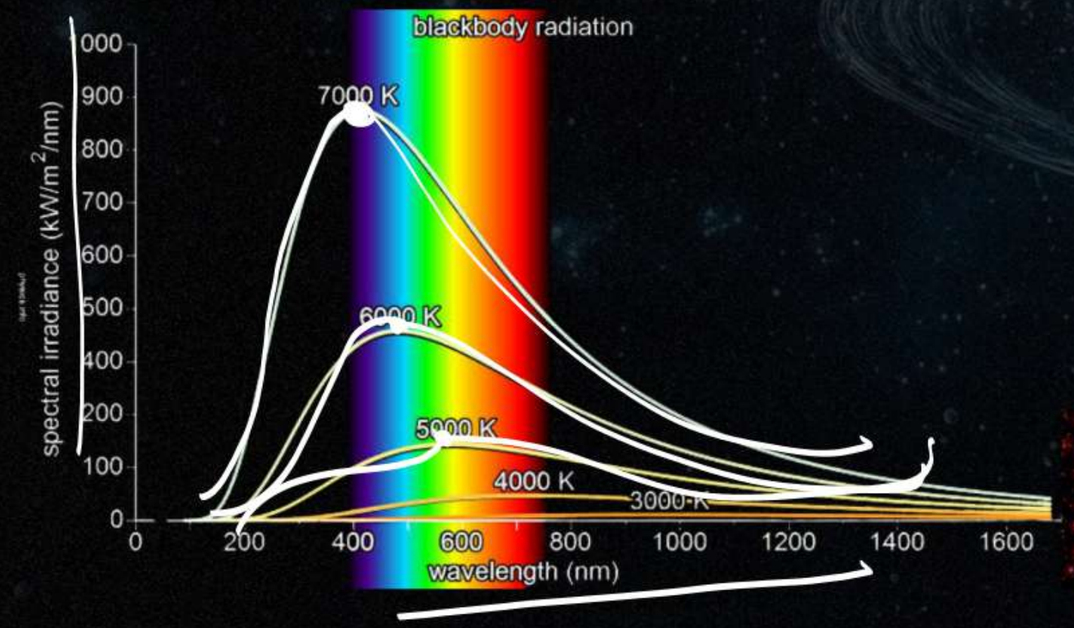


$T_1$ Red

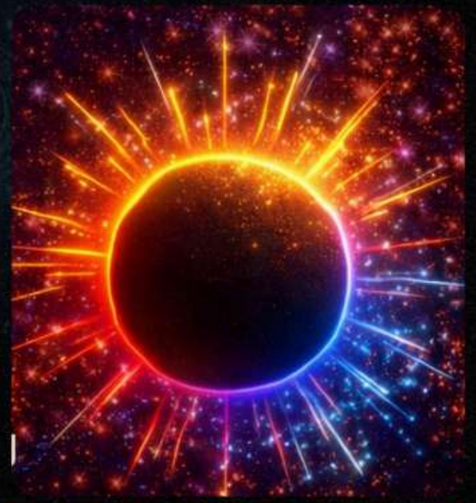
$T_2$ Blue

# Black Body Radiation

- An ideal body, which emits and absorbs radiations of all frequencies uniformly, is called a black body and the radiation emitted by such a body is called black body radiation.
- In practice, no such body exists. Carbon black approximates fairly closely to black body.
- The different colours that come out of the black body depends on it's temperature.



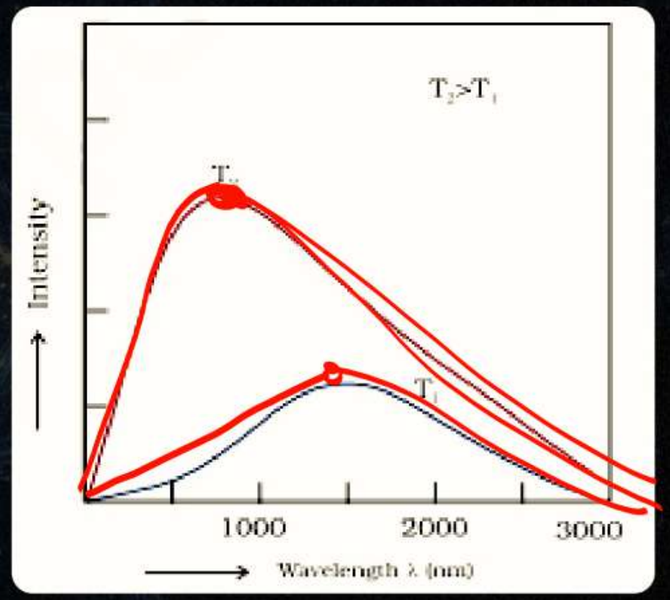




# Black Body Radiation

- Blackbody spectrum depends only on the temperature of the object, and not on what it is made of.
- As the temperature of an object increases, it emits more blackbody energy at all wavelengths.
- As the temperature of an object increases, the peak wavelength of the blackbody spectrum becomes shorter (blue).
- The blackbody spectrum always becomes small at the left-hand side (the short wavelength, high frequency side).



# Particle Nature of Electromagnetic Radiation: Planck's Quantum Theory – E = hν → frequency.

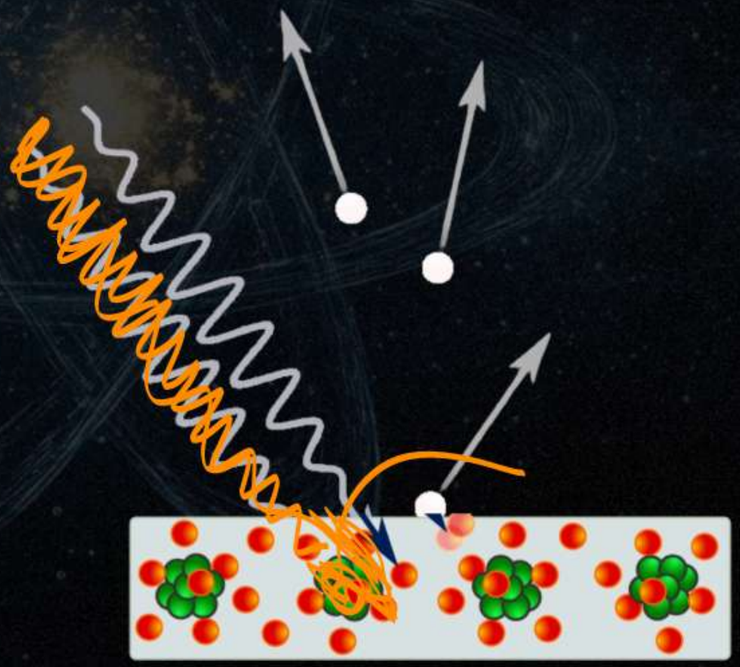
energy Planck const.

h = 6.626 × 10⁻³⁴ JS

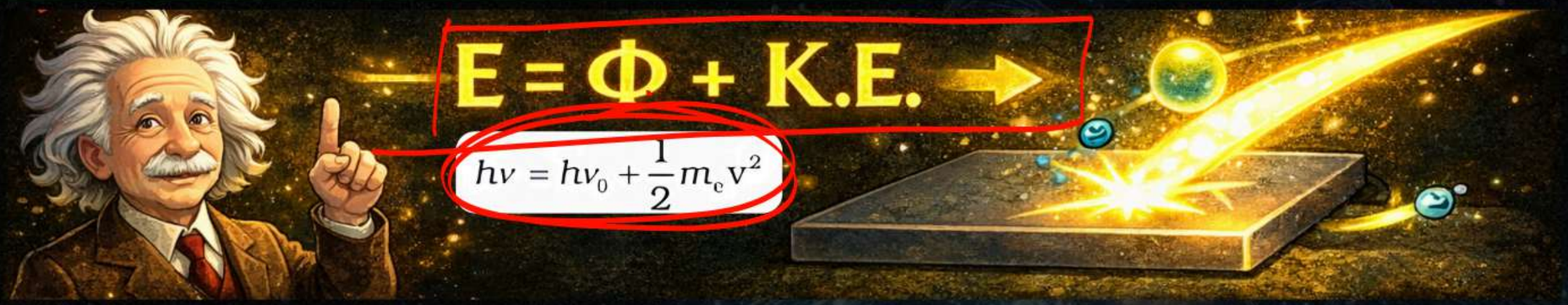
- The radiant energy or electromagnetic radiation is emitted or absorbed discontinuously in the form of small discrete packets of energy (singular: quantum and plural: quanta) and not continuously.
- The nature of emission of radiation from hot bodies (black-body radiation) ejection of electrons from metal surface when radiation strikes it (photoelectric effect) variation of heat capacity of solids as a function of temperature.
- Line spectra of atoms with special reference to hydrogen. These phenomena indicate that the system can take energy only in discrete amounts.
- All possible energies cannot be taken up or radiated.

# Photoelectric Effect

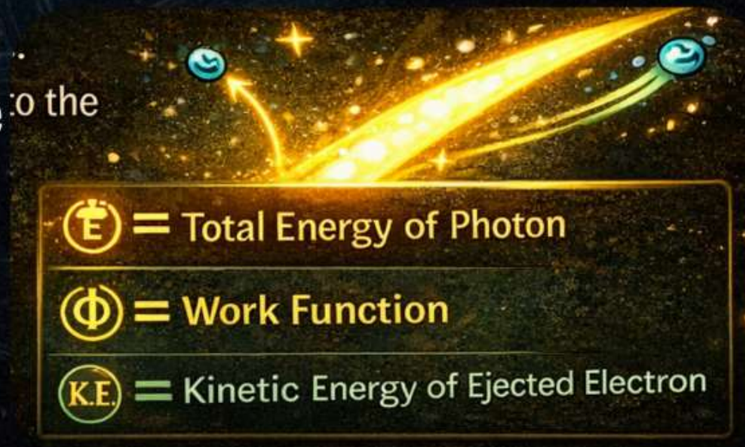
- In 1887, H. Hertz performed a very interesting experiment in which electrons (or electric current) were ejected when certain metals
  - (for example potassium, rubidium, caesium etc.) were exposed to a beam of light.
- The phenomenon is called Photoelectric Effect.
- The electrons are ejected from the metal surface as soon as the beam of light strikes the surface, i.e., there is no time lag between the striking of light beam and the ejection of electrons from the metal surface.
- The number of electrons ejected is proportional to the intensity or brightness of light.



# Einstein's Photoelectric Equation




- The total energy of the incoming photon (E) must be equal to the minimum energy required to eject the electron from the metal (Φ) + kinetic energy of the ejected electron (K.E.).
- In the photoelectric effect, the work function (Φ) is the minimum amount of energy ($h\nu_0$) needed to eject an electron from a metal surface. It represents the strength of the bond holding electrons to the metal's lattice.
- This energy is specific to each material and is directly related to the threshold frequency of light required to initiate the photoelectric effect.

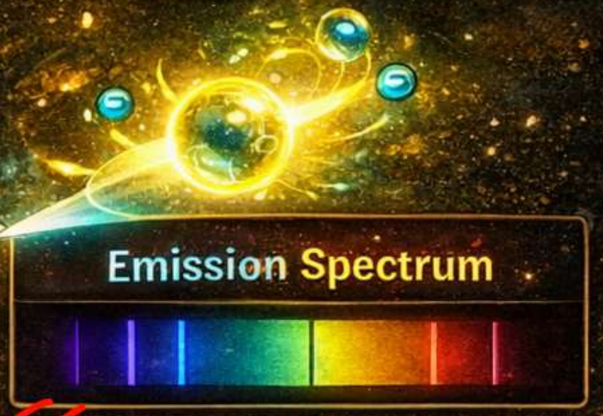


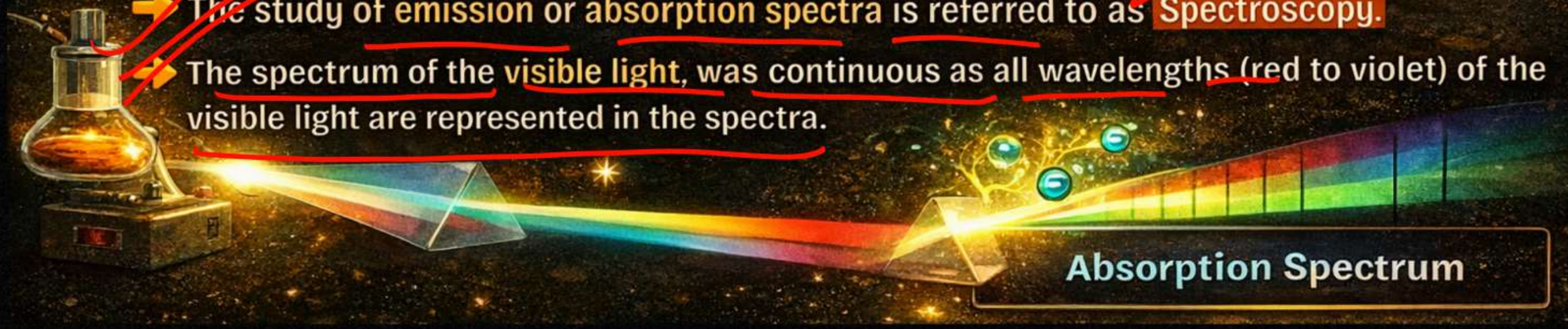
# Emission & Absorption Spectra



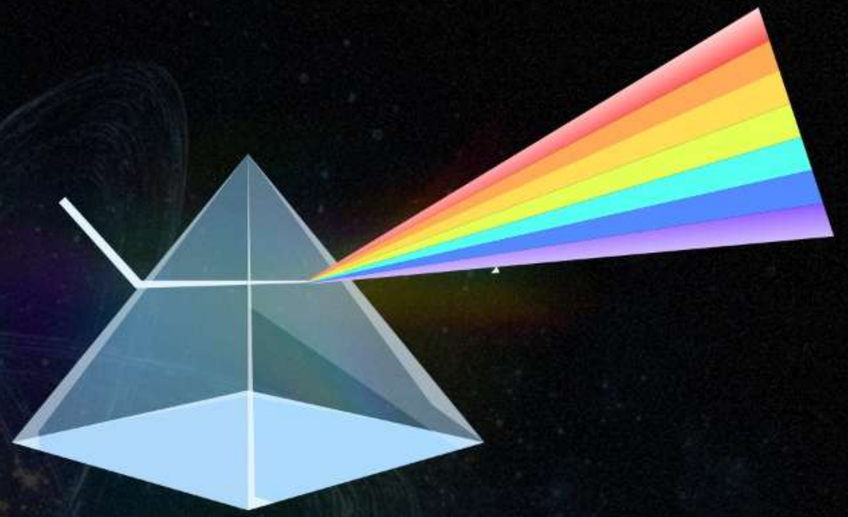
- The spectrum of radiation emitted by a substance that has absorbed energy is called an **Emission Spectrum.**
- Atoms, molecules or ions that have **absorbed radiation** are said to be "**excited**".
- To produce an **emission spectrum, energy** is supplied to a sample by heating it or irradiating it and **the wavelength (or frequency)** of the radiation emitted, as the sample gives up the absorbed energy, is recorded.
- The study of **emission** or **absorption spectra** is referred to as **Spectroscopy.**
- The spectrum of the **visible light**, was continuous as all wavelengths (red to violet) of the visible light are represented in the spectra.



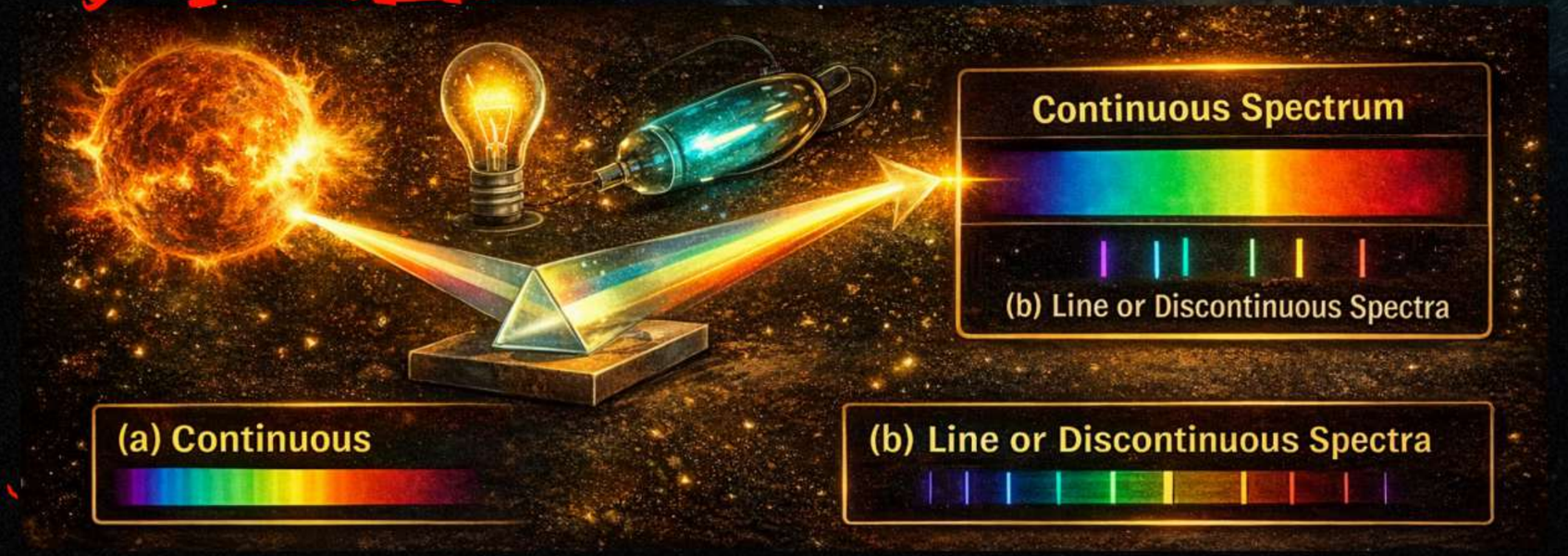



# Emission Spectra



- When radiation emitted from a source like Sun, hot bulb or from a gas/matter whose electrons are excited by electrical or heat energy is passed directly through a prism and received on a photographic plate is called emission spectrum.
- Depending upon the source of radiation, emission spectra is classified into (a) Continuous (b) Line or Discontinuous Spectra

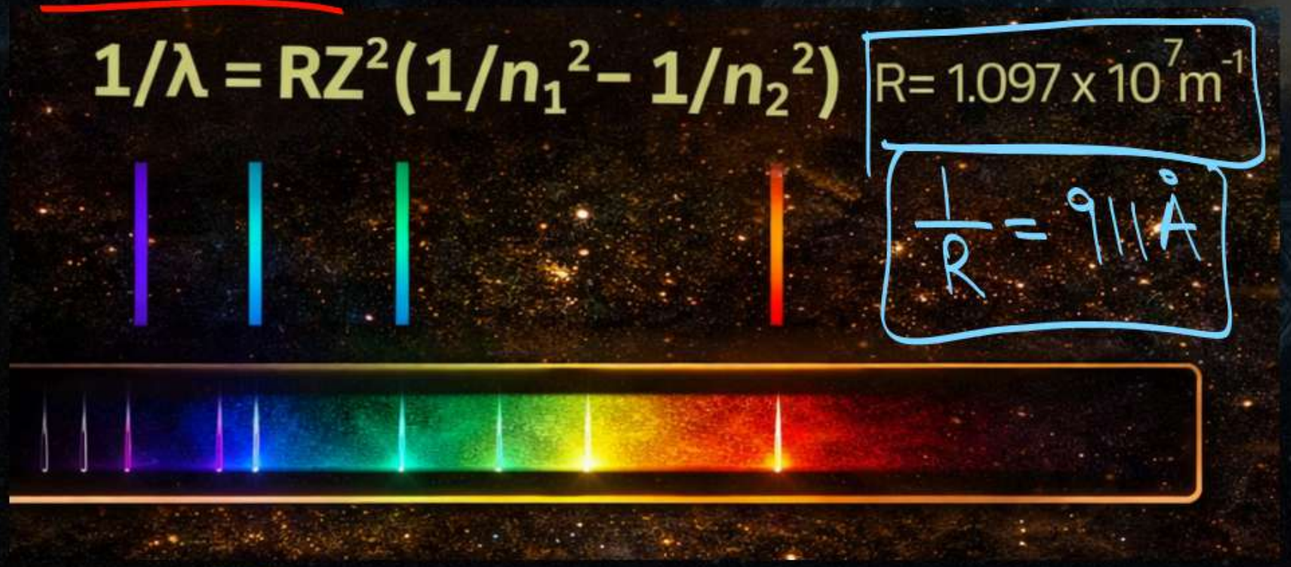


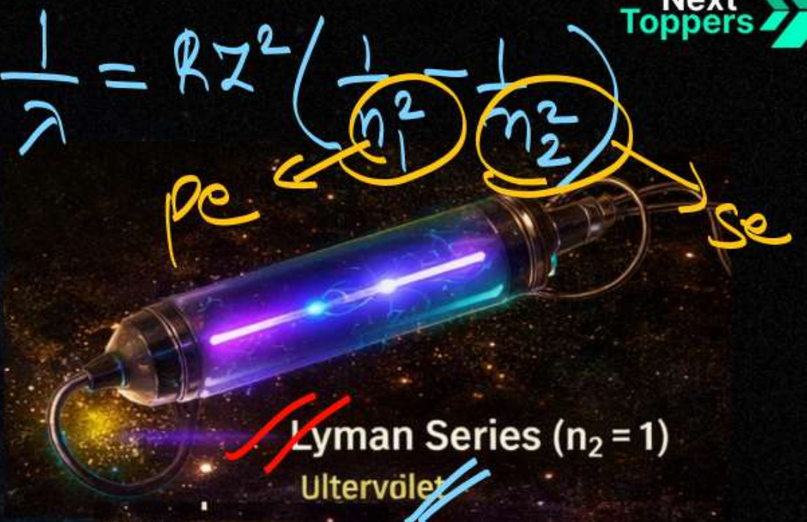
# Line Spectrum of Hydrogen

$\bar{\nu} = \frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

pe → $n_1^2$ $n_2^2$ ← se

- When an electric discharge is passed through gaseous hydrogen, the $H_2$ molecules dissociate and the energetically excited hydrogen atoms produced emit electromagnetic radiation of discrete frequencies.
- The hydrogen spectrum consists of several series of lines named after their discoverers.



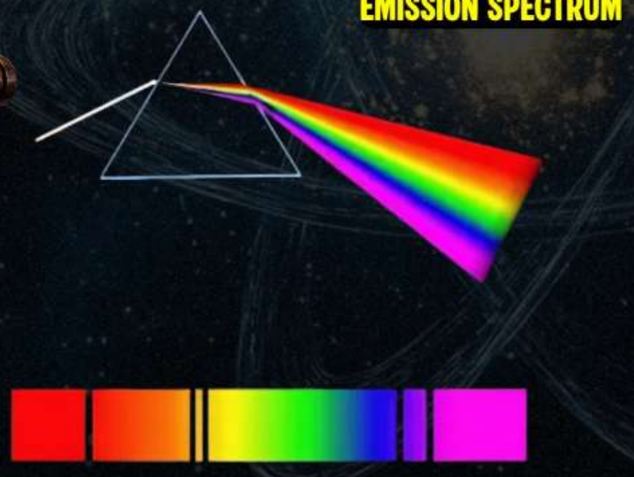



- Lyman Series ($n_2 = 1$) — Ultervölet
- Balmer Series ($n_2 = 2$) — Visible
- Paschen Series ($n_2 = 3$) — Infrared
- Brachett Series ($n_2 = 4$) — Infrared
- Pfund Series ($n_2 = 5$) — Infrared
- Humphrey Series ($n_2 = 6$) — Infrared

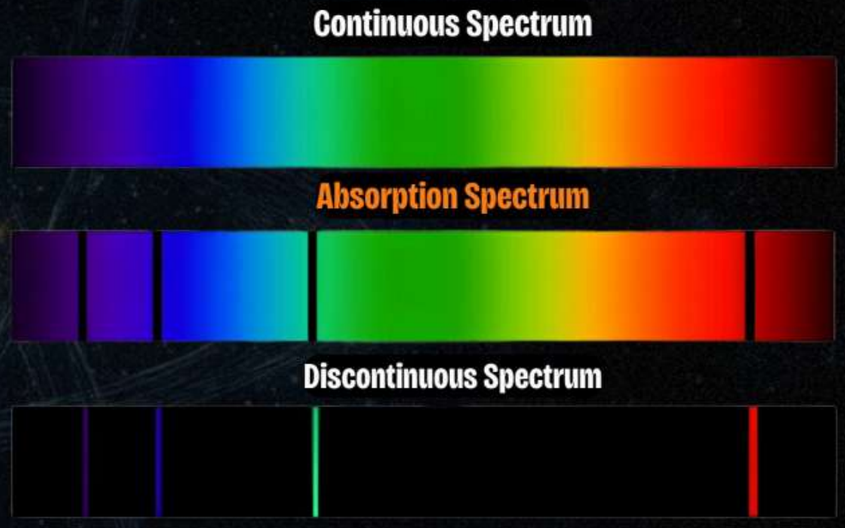
# Absorption Spectrum

- When white light is first passed the solution or vapours of a chemical substance and then analysed in(spectroscope some dark lines are observed at the same place where coloured lines were obtained in the emission spectrum.



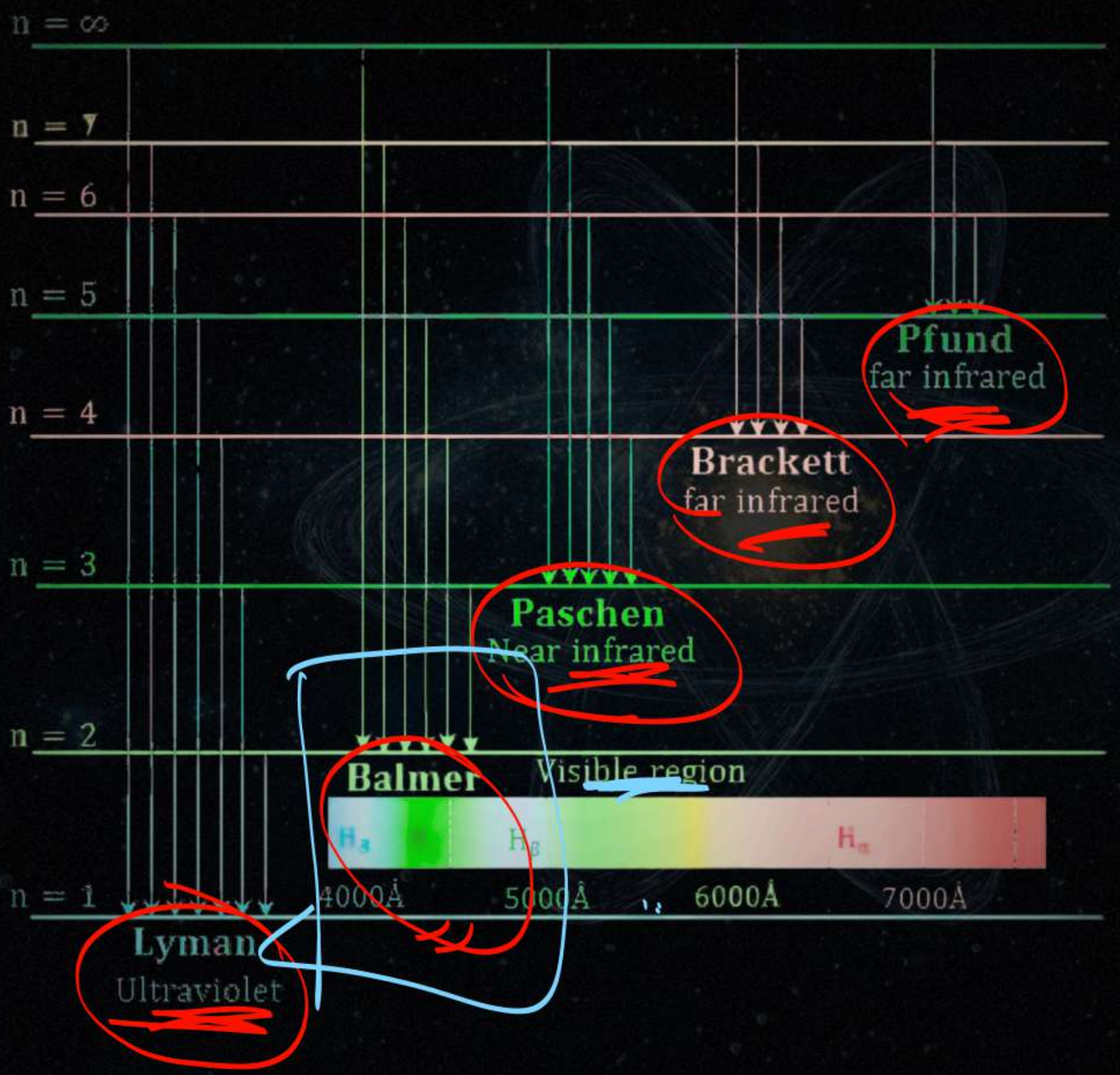





BACKLOG MARATHON SERIES

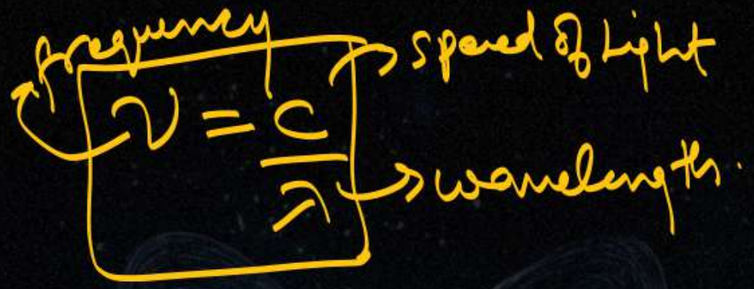
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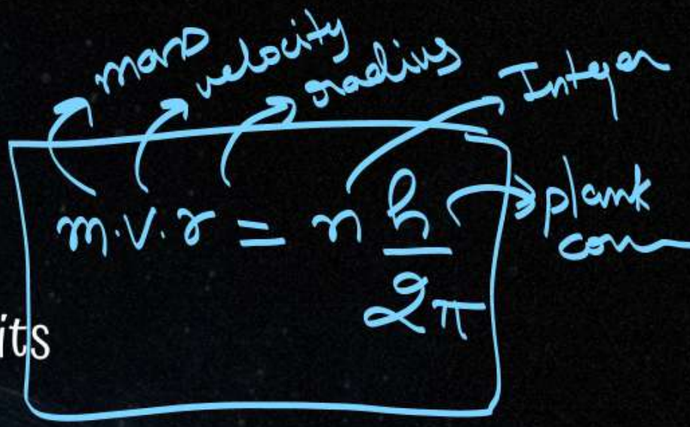


BACKLOG MARATHON SERIES

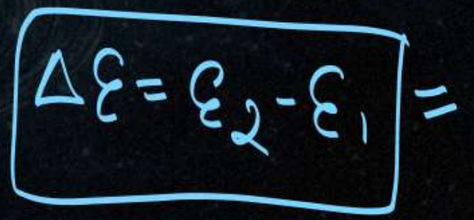
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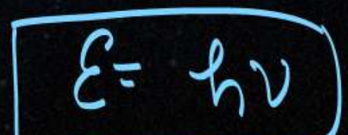
# Bohr's Atomic Model

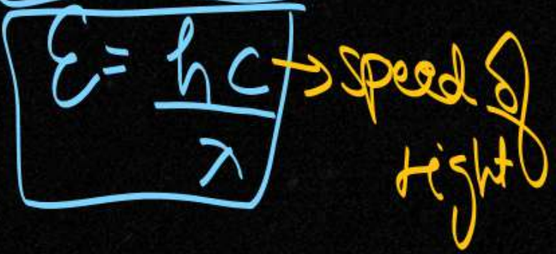


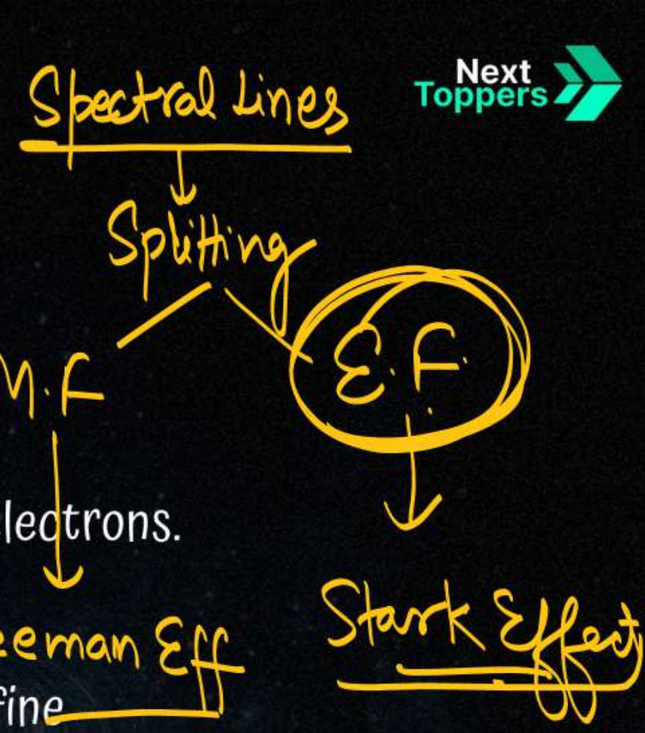

- Atom has a centre called nucleus . Electrons revolve only in fixed circular orbits with fixed energy & fixed velocity .
- **Quantisation condition-** Electrons revolve only in those circular orbits for which their angular momentum is an integral multiple of h/2π.
- **Angular momentum is defined as:** The property of any rotating object given by moment of inertia times angular velocity.
- Energy is emitted or absorbed only when an electron Jumps from higher energy level to lower energy level and vice-versa.
- del E = E2 – E1 = hv = hc / lamda is known as Bohr's frequecy rule.
- The most stable state of an atom is its ground state or normal state.







# Drawbacks of Bohr's Atomic Model




## 1. Limited Applicability:

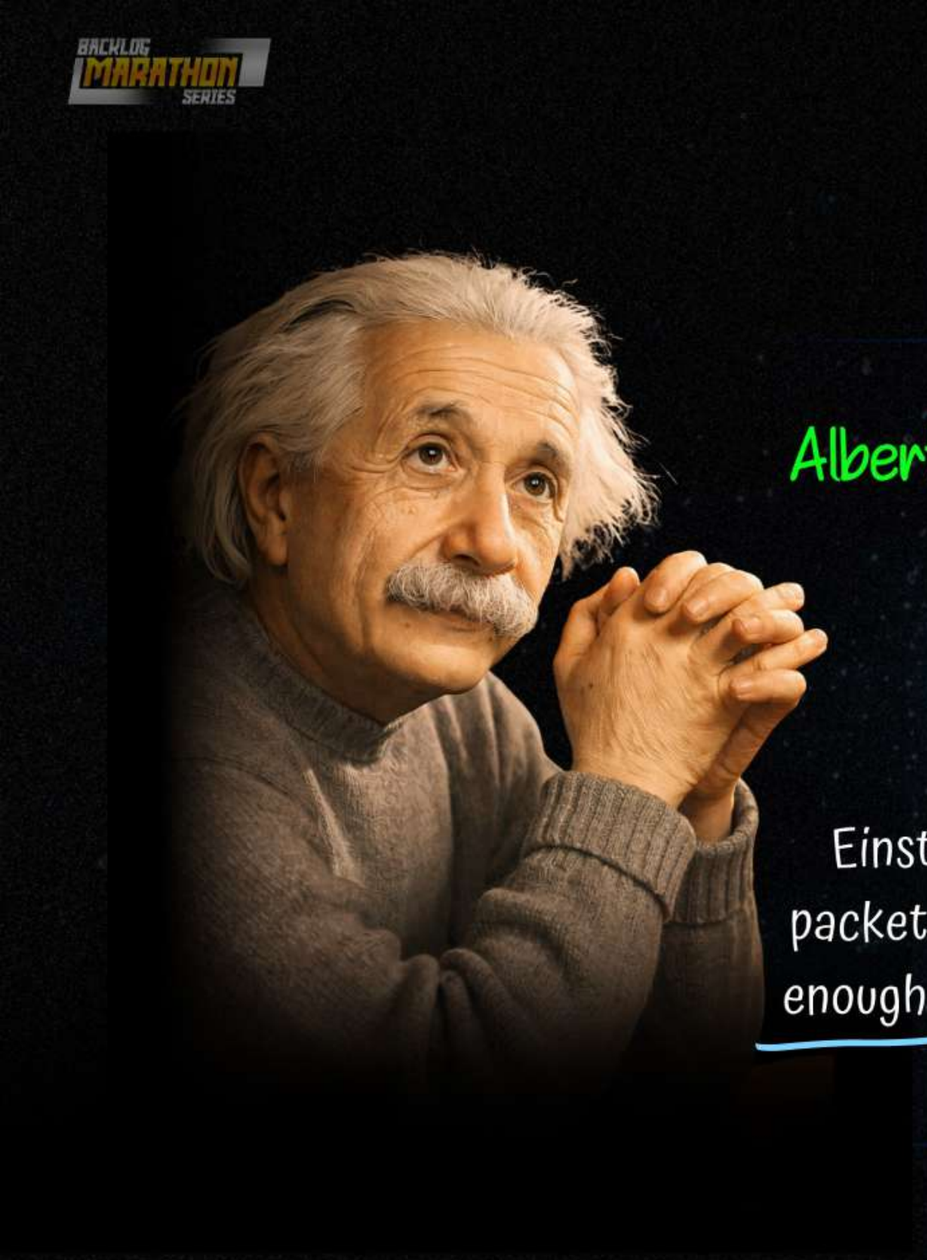
- Bohr's model is primarily successful in explaining the hydrogen atom's spectrum.
- It struggles to accurately predict the spectra of more complex atoms with multiple electrons.

## 2. Inability to Explain Fine Structure:

- Bohr's model predicts single spectral lines, but in reality, many spectral lines exhibit a fine structure (splitting into multiple lines).
- This fine structure arises from relativistic effects and spin-orbit coupling, which Bohr's model doesn't account for.

## 3. Failure to Explain Zeeman and Stark Effects:

- The Zeeman effect describes the splitting of spectral lines in a magnetic field.
- The Stark effect describes the splitting of spectral lines in an electric field.
- Bohr's model cannot explain either of these effects.






## Albert Einstein explained the photoelectric effect in 1905:

### Explanation

Einstein's theory was based on the idea that light is made up of tiny packets of energy called photons. He explained that when photons with enough energy hit a metal surface, they can transfer their energy to the electrons in the metal, causing them to be ejected.

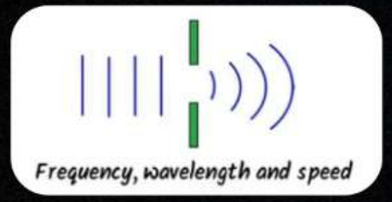
BACKLOG MARATHON SERIES

Next Toppers

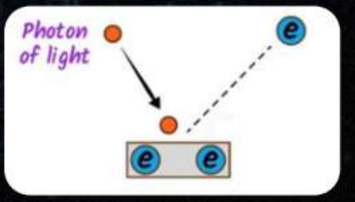
# Wave Particle Nature of Light



Wave diffraction



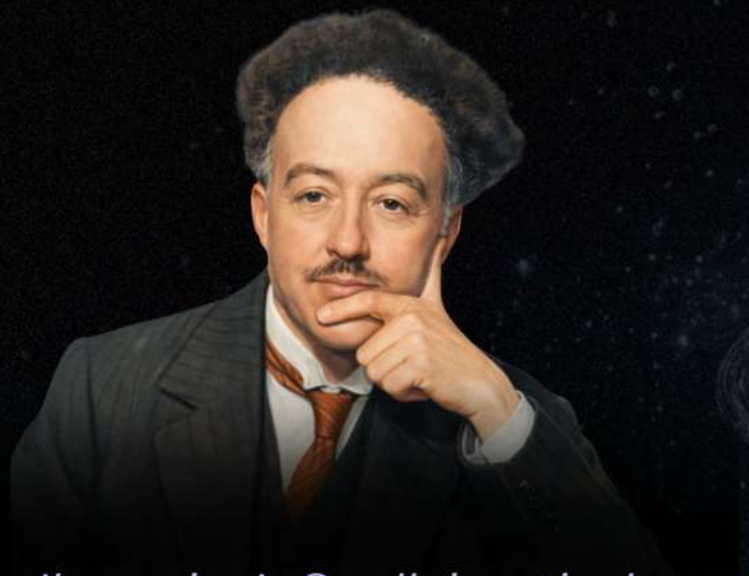

Particle photoelectric effect




**All electromagnetic waves behave like a wave & particle**

# De-Broglie Hypothesis

$p = mv$



According to the de-Broglie hypothesis,a moving material particle behaves like a wave at times and like a particle at other times.
Every moving material particle is connected with a wave.

- De Broglie's Hypothesis says that Matter consists of both the particle nature as well as wave nature.
- De Broglie wavelengthis given as,

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

$p = \sqrt{2 K.E\, m_e}$

- From the above relation, it can be said that the wavelength of the matter is inversely proportional to the magnitude of the particle's linear momentum.

Light → wave → particle

Electron → particle → wave.

# De-Broglie Hypothesis

$$\lambda = \frac{h}{m \cdot v}$$

- This relation is applicable to both microscopic and macroscopic particles.
- The de Broglie equation is one of the equations that is commonly used to define the wave properties of matter.
- Electromagnetic radiation exhibits the dual nature of a particle (having a momentum) and wave (expressed in frequency, and wavelength).



Davisson and L.H. Germer performed an experiment known as the Davisson Germer experiment to explain the wave nature of electrons through electron diffraction.

**Que:-** Why we cannot observe these waves?

**Ans:-** Because the wavelength of their motion is extremely small



# Heisenberg's Uncertainty Principle



Heisenberg's uncertainty principle states that it is impossible to measure or calculate exactly both the position and the momentum of an object.

- This principle is based on the wave-particle duality of matter.
- Although Heisenberg's uncertainty principle can be ignored in the macroscopic world (the uncertainties in the position and velocity of objects with relatively large masses are negligible), it holds significant value in the quantum world.
- Since atoms and subatomic particles have very small masses, any increase in the accuracy of their positions will be accompanied by an increase in the uncertainty associated with their velocities.

Uncertainity in momentum

$$\Delta x \Delta p \geq \frac{h}{4\pi}$$

Uncertainity in position


Bhaiya kaha ana hai ???
India ✓
Mumbai ✓
Indira Nagar ✓
Probability.

# Location of an Electron

- A large number of orbitals are possible in an atom.
- Qualitatively these orbitals can be distinguished by their size, shape and orientation.
- An orbital of smaller size means there is more chance of finding the electron near the nucleus.
- Similarly shape and orientation mean that there is more probability of finding the electron along certain directions than along others.

**Atomic orbitals are precisely distinguished by what are known as quantum numbers.**

- Principal quantum number (n) → shell → size
- Azimuthal or Orbital or Subsidary quantum number (l) → subshell, shape
- Magnetic quantum number (m) → orbital → orientation
- Spin quantum number (s) → direction

# Principal Quantum number (n)

It tells us about the main shell in which electron resides. It also gives an idea about the energy of shell and average distance of the electron from the nucleus (Size).

**Value of n = any integer.**

**n= 1,2,3,..... for shell K, L, M, ..**

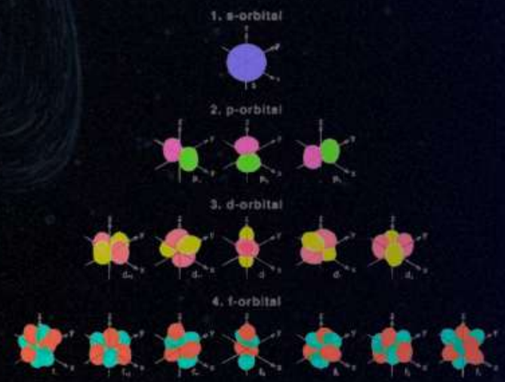
# Azimuthal Quantum number (l)

- It tells about the number of subshells (s. p, d, f) in any main shell and shapes of subshells.
- It defines the three-dimensional shape of the orbital. For a given value of n, l can have n values ranging from 0 to n - 1, that is, for a given value of n, the possible value of l are : *l* = 0, 1, 2, .... (n-1) .

**For example,** when n = 1, value of l is only 0. For n = 2, the possible value of l can be 0 and 1. For n = 3, the possible l values are 0, 1 and 2.

Azimuthal Quantum numbers also tells about the shape of the orbital .

**NOTE** Orbital is the 3-D discription where the probability of finding an electron is maximum



# Magnetic Quantum number (m)

- It tells about the number of orbitals and orientation of each subshell.
- Value of m = -l to +l including zero.
- Number of orbitals in each subshell = (2l + 1)
- Number of orbitals in main energy level = n2

*d* sublevel, *l* = 2



$m_l$ = -2, -1, 0, +1, +2

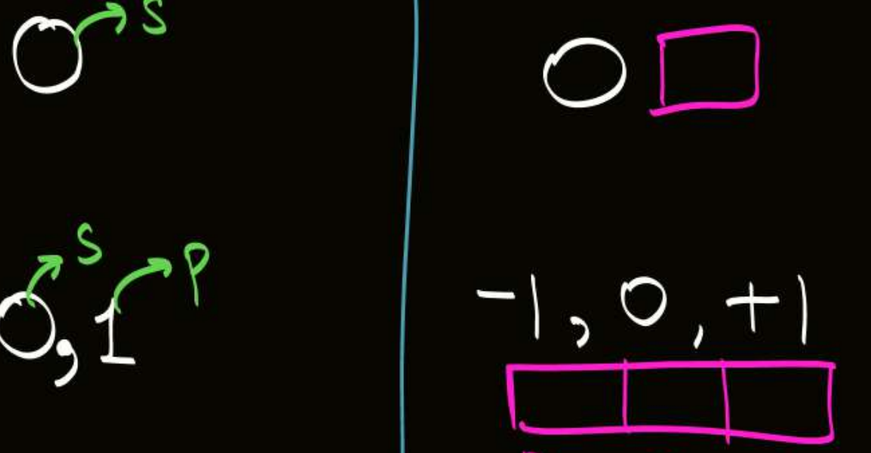
# Spin Quantum number (S)

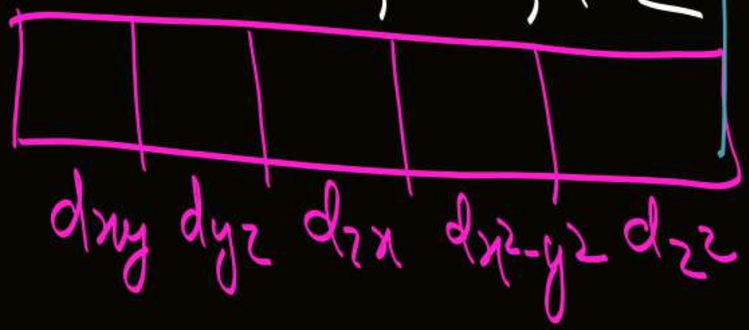
- It indicates the direction of spinning of electron, i.e., clockwise or anti-clockwise.
- Shows how the electron can rotate about its own axis .
- Maximum number of electrons in main energy level = $2n^2$.

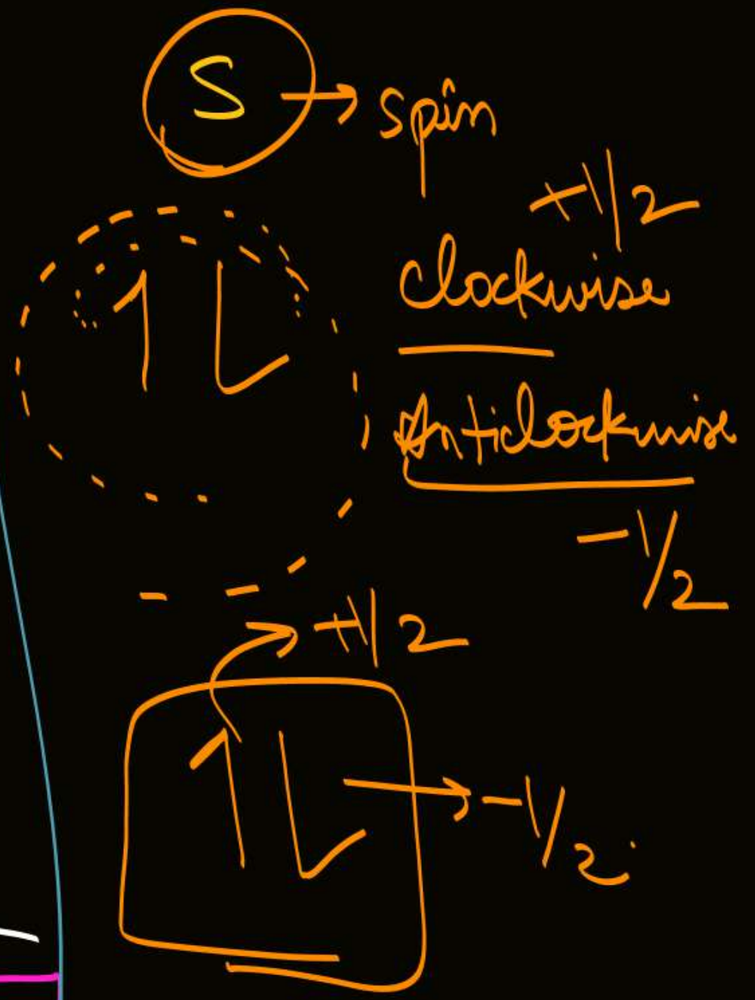



*ms* = + 1/2 *ms* = - 1/2

| $n$ | $0 \le l < n-1$ | $-l < m < +l$ |
|---|---|---|
| 1 | 0 (s) | 0 |
| 2 | 0 (s), 1 (p) | −1, 0, +1 |
| 3 | 0 (s), 1 (p), 2 (d) | −2, −1, 0, +1, +2 |



1 orbital → max ↓ 2e⁻

Subshell. ←

0 → s → spherical

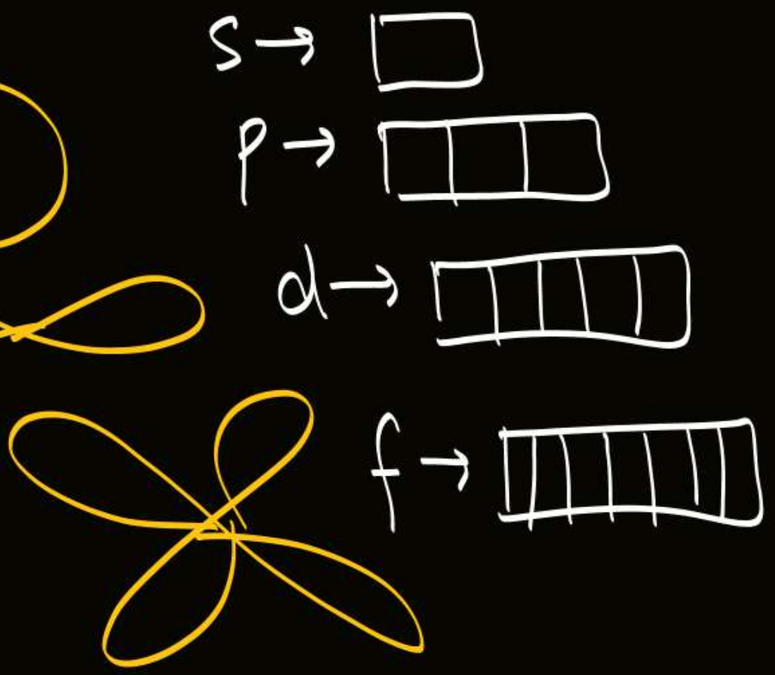
1 → p → dumbell

2 → d → double dumbell

3 → f → complex.

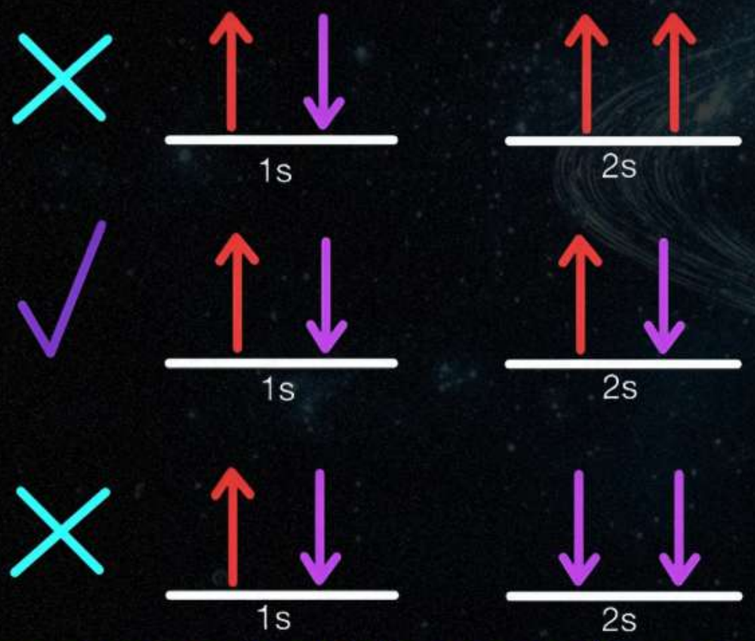
inside ↓

orbital → 3-D → Prob. of finding an electron is max.

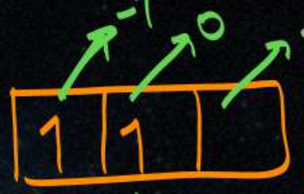


# Pauli Exclusion Principle

- It states, no two electrons in an atom can have identical set of four quantum numbers.
- Only two electrons may exist in the same orbital and these electrons must have opposite spin".




$C \rightarrow 1s^2 2s^2 2p^2$




$n = 2^{nd}$ $n = 2^{nd}$

$l = 1$ $l = 1$

$m = -1$ $m = 0$

$s \rightarrow +1/2$ $m = +1/2$

**The maximum number of electrons in s subshell is 2, p subshell is 6 d subshell is 10 and f subshell is 14.**

# Aufbau Principle:

$n+l \propto$ Energy.

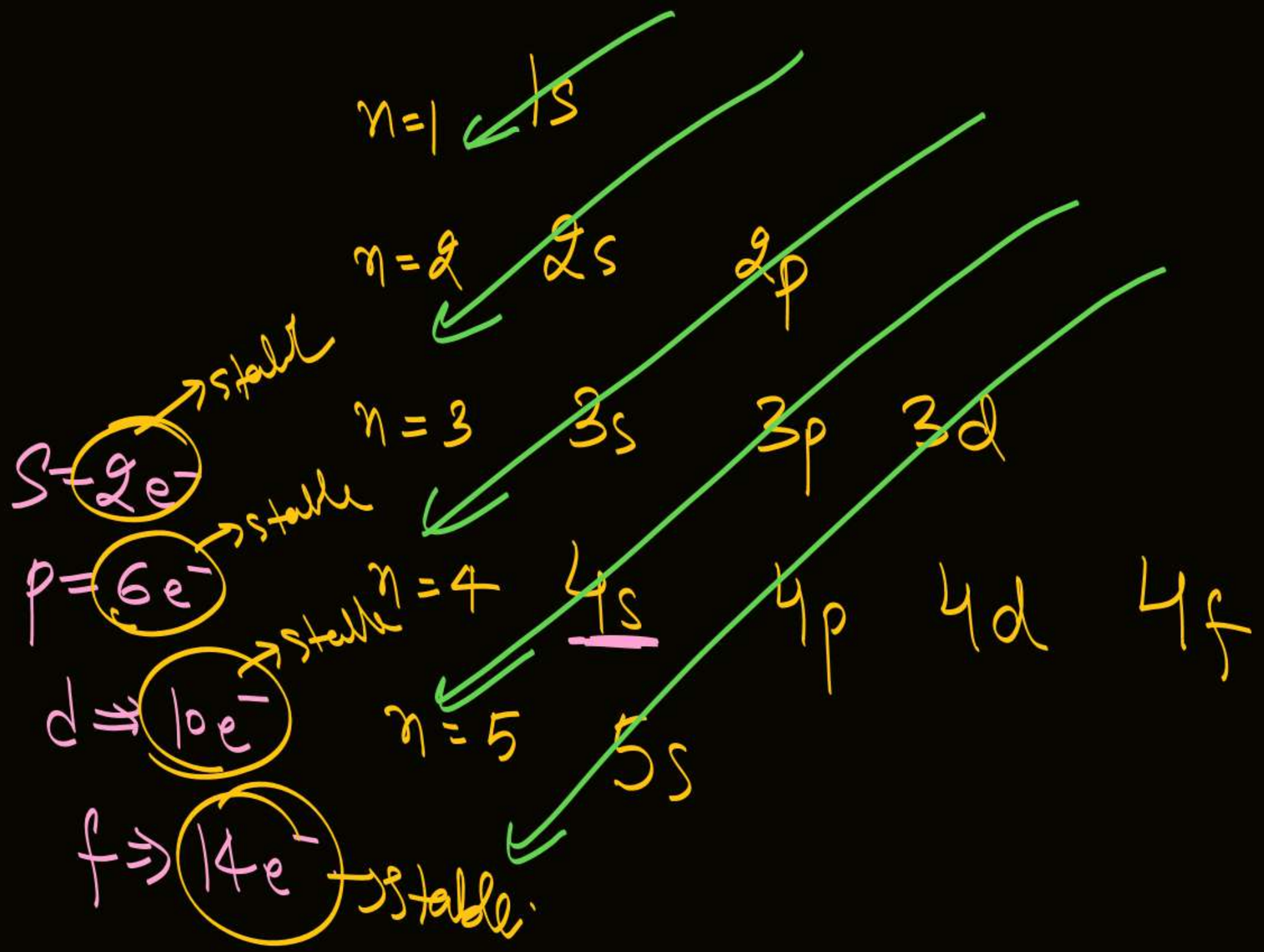
**Afbau is a German word which means structure or constructuction.**

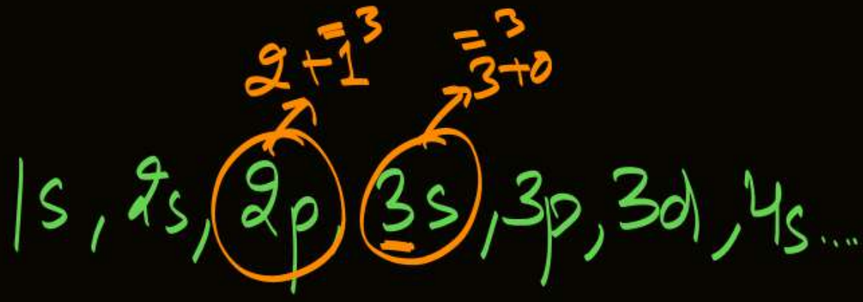
- Aufbau Principle states that in the ground state of an atom, the electrons fill atomic orbitals of the lowest available energy levels before occupying higher levels i.e. orbitals are filled in the increasing order of their energies.
- The rule may be used to identify the sequence in which the energy of orbitals grows, where the sum of the main and azimuthal quantum numbers defines the energy level of the orbital.
- Lower values indicate lower orbital energy. If two orbitals have equivalent values, the orbital with the lower value is said to have lower energy.
- If value of (n + l) is same for any two subshell, then the deciding factor will be the value of n. Greater the value of n, greater will be its energy.

**The orbitals are filled with electrons in the following order:**

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p , 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p. .

**and so on.**


n=1 1s
n=2 2s 2p
n=3 3s 3p 3d
n=4 4s 4p 4d 4f
n=5 5s
s⇒2e⁻ stable
p⇒6e⁻ stable
d⇒10e⁻ stable
f⇒14e⁻ stable.


2+1=3
3=3+0
1s, 2s, 2p, 3s, 3p, 3d, 4s....


same → jiski n ki value badi uski zyada Energy.
n+l
choti
badi
e⁻ pehle jayega.
e⁻ baad jayega.

$_{3}Li \rightarrow 1s^2 2s^1$

$_{8}O \rightarrow 1s^2 2s^2 2p^4$

$_{20}Ca \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

$_{23}V \rightarrow \boxed{1s^2 2s^2 2p^6 3s^2 3p^6} 4s^2 3d^3$

$\boxed{[Ar]\, 4s^2\, 3d^3}$

## Exceptions

$_{29}Cu$ Copper

$[Ar]\, 4s^2\, 3d^9$

$\boxed{[Ar]\, 4s^1\, 3d^{10}}$

$_{24}Cr$ Chromium.

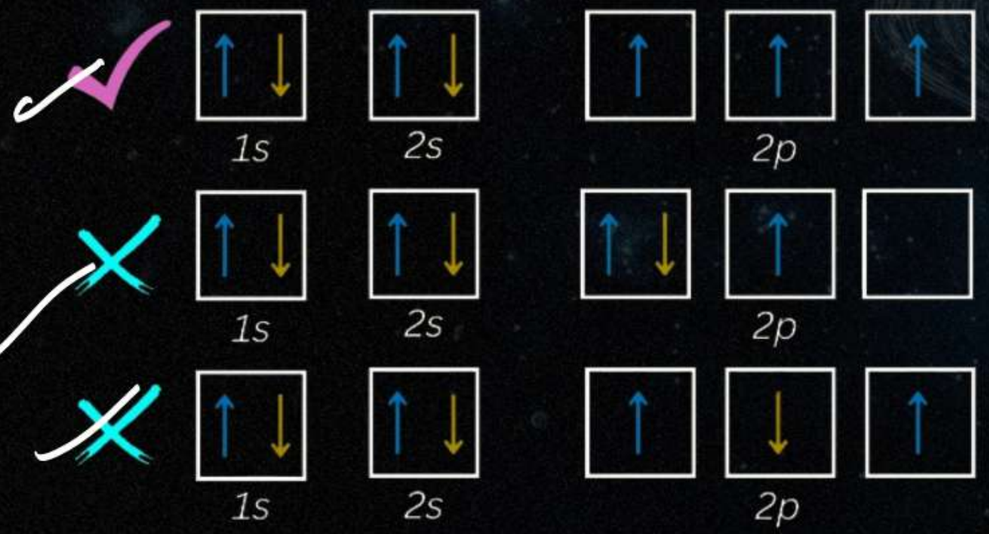
$[Ar]\, 4s^2\, 3d^4$

$\boxed{[Ar]\, 4s^1\, 3d^5}$.

# Hund's Rule

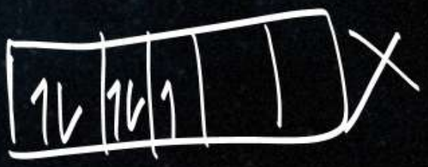
e → filling → half filled → same spin
pairing → opp spin.

- Hunds Rule of Maximum Multiplicity rule states that for a given electron configuration, the term with maximum multiplicity falls lowest in energy.
- According to this rule electron pairing in p, d and f orbitals cannot occur until each orbital of a given subshell contains one electron each or is singly occupied.









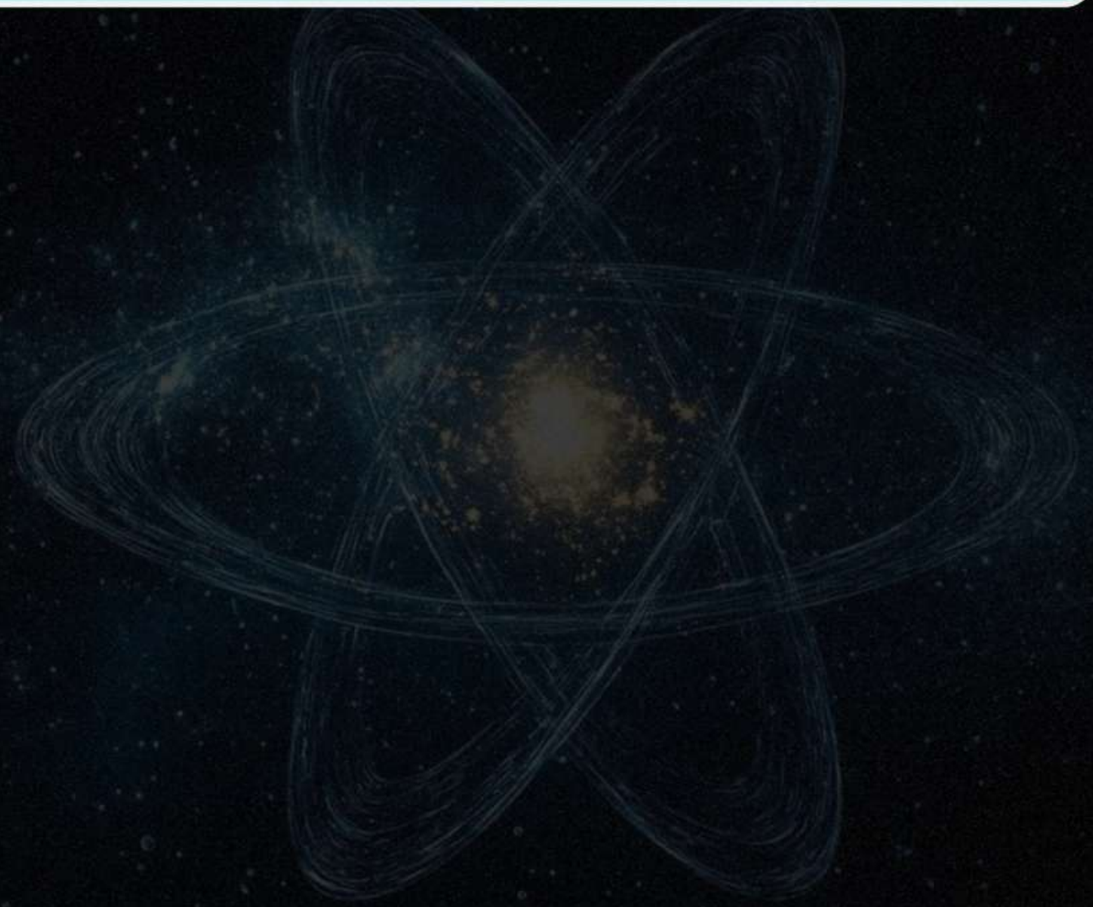




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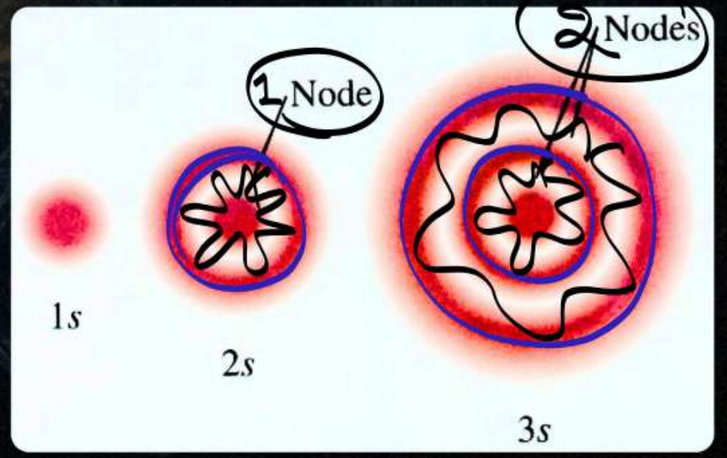
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# Electronic Configuration upto 30 Elements



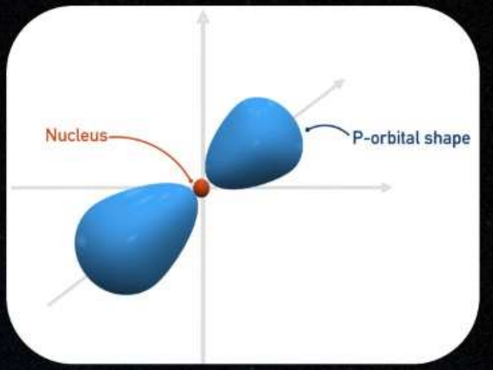
# Shape of Orbital:

- Spherical in shape. It's nodes are also known as radial nodes.
- $[\Psi]^2$ gives the electron probability density of the electron at any point.
- The probability of 1s electron is found to be maximum near the nucleus and decreases as the distance from the nucleus increases.
- The intermediate region (a spherical shell) where the probability is zero is called node.
- Nodal Surface - Radial Node whereas Nodal Plane - Angular Node
- No. of nodes (spherical nodes) = n-l -1
- In general, any ns orbital has (n-l-1) nodes. (spherical nodes only)
- Electron density is equal in all direction (nondirectional).



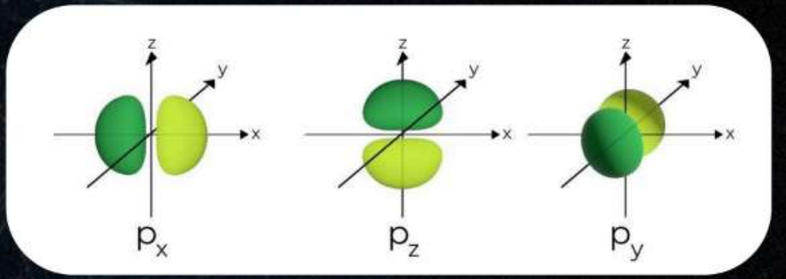
# Shape of Orbital:






## Shape of p-Orbital

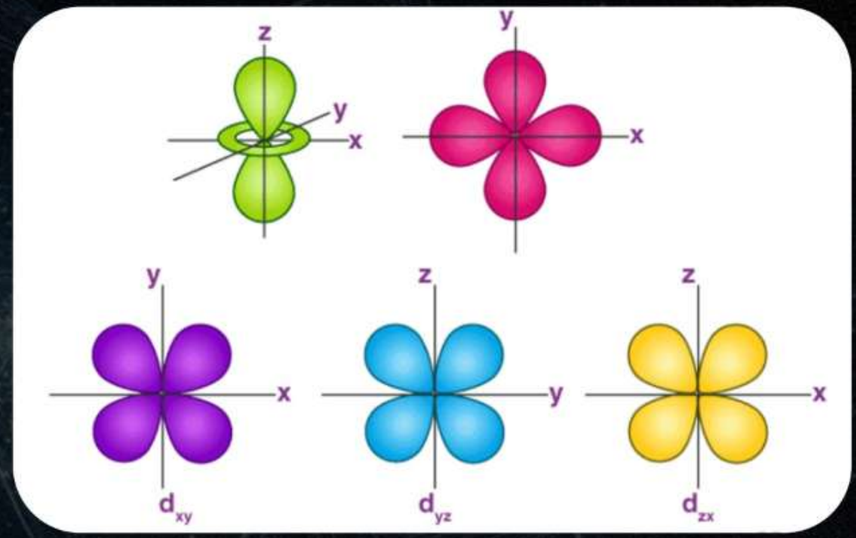
- The probability of finding the p-electrons is maximum in two lobes on the opposite sides of the nucleus. This gives rise to a "dumb-bell shape" for the p-orbitals.
- A plane passing through the nucleus on which the probability of finding the electron is almost zero. This is called nodal plane .ual energies and are called degenerate.
- p-orbitals have three different orientations (x, y, z). These are designated as px, py and pz.
- Unlike s-orbitals, p-orbitals have directional Characteristics.
- The three p-orbitals belonging to a particular energy shell have equal energies and are called degenerate.

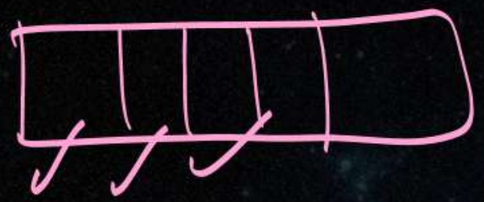


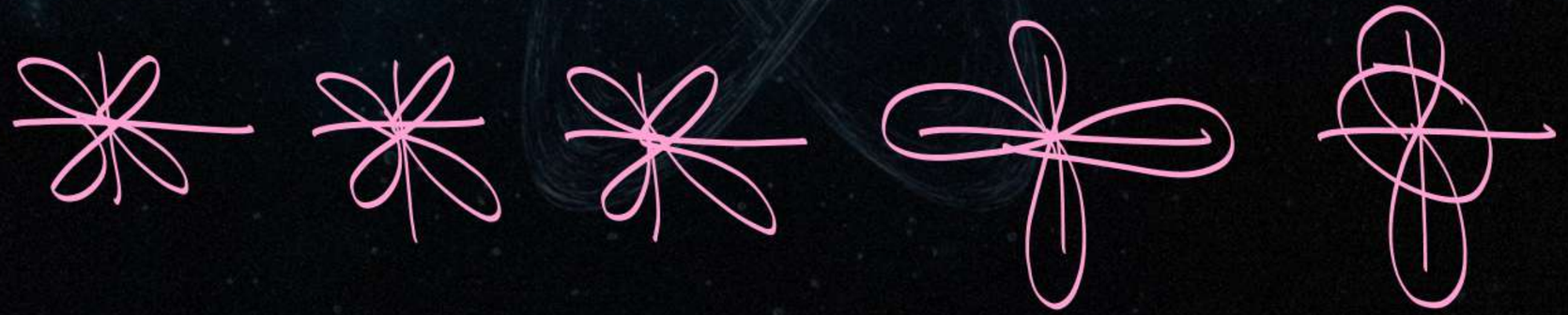
# Shape of Orbital:

## Shape of d-Orbital

- Double dumbbell shape (except $d_{Z^2}$) .
- $d_{Z^2}$ doughnut shaped or peanut shell shaped .
- dxy , dyz , dzx are non axial.





# Nodes

→ Radial → s orbital → non-directional

principal Q.N.
Azimuthal Q.N.

$n - l - 1$ = node

→ Angular → p orbital, d orbital } directional → node = $l$

3p → nodes?

Angular node → $l \rightarrow 1$

4s → Radial → $n - l - 1$

$l \rightarrow 0$

$4 - 0 - 1$

3

Using s, p, d, f notations, describe the orbital with the following quantum numbers
(a) n = 2, l = 1, (b) n = 4, l = 0, (c) n = 5, l = 3, (d) n = 3, l = 2.

2p

4s


Apne
itne acche se
pdha
vps a jana
21:55
isliye meri taraf se
15 minute ka
break lelo
Keep Going!
Proud of you
You Did Great!
Break Time
Hard Work Today
Better Tomorrow
Future is Bright
Notes

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# Class 11th | Chemistry

## Classification of Elements & Periodicity in Properties

## Chapter 3

6 Marks

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# Why do we need to Classify Elements

- Systematic Study
- Predicting Properties
- Understanding Relationships
- Facilitating Research and Discovery
- Comparing Properties



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# Classification of Elements & Periodicity in Properties

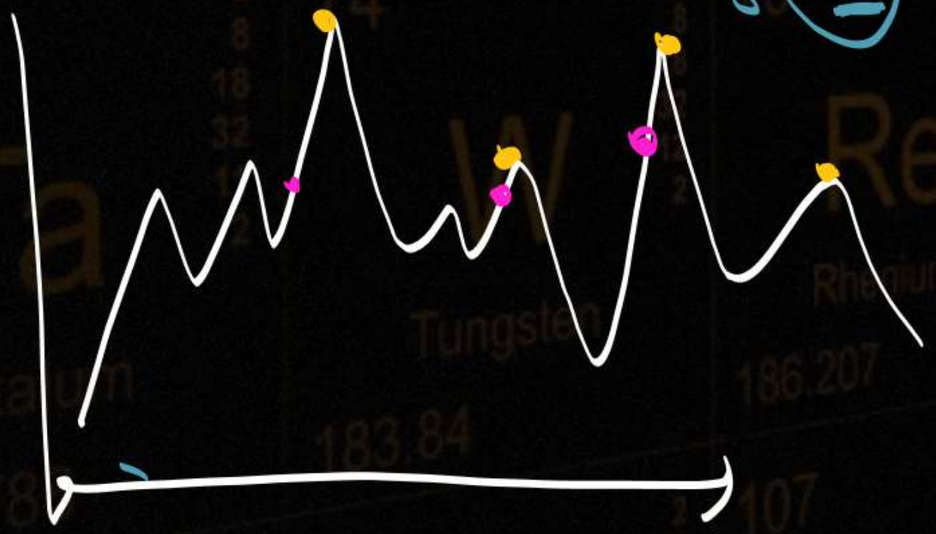
- Lavoisier's Classification
- Döbereiner Triads
- Newlands' Law of Octaves
- Lothar meyer periodic table

Metal
Non Metal

Li 7 | Na 23 | K 39

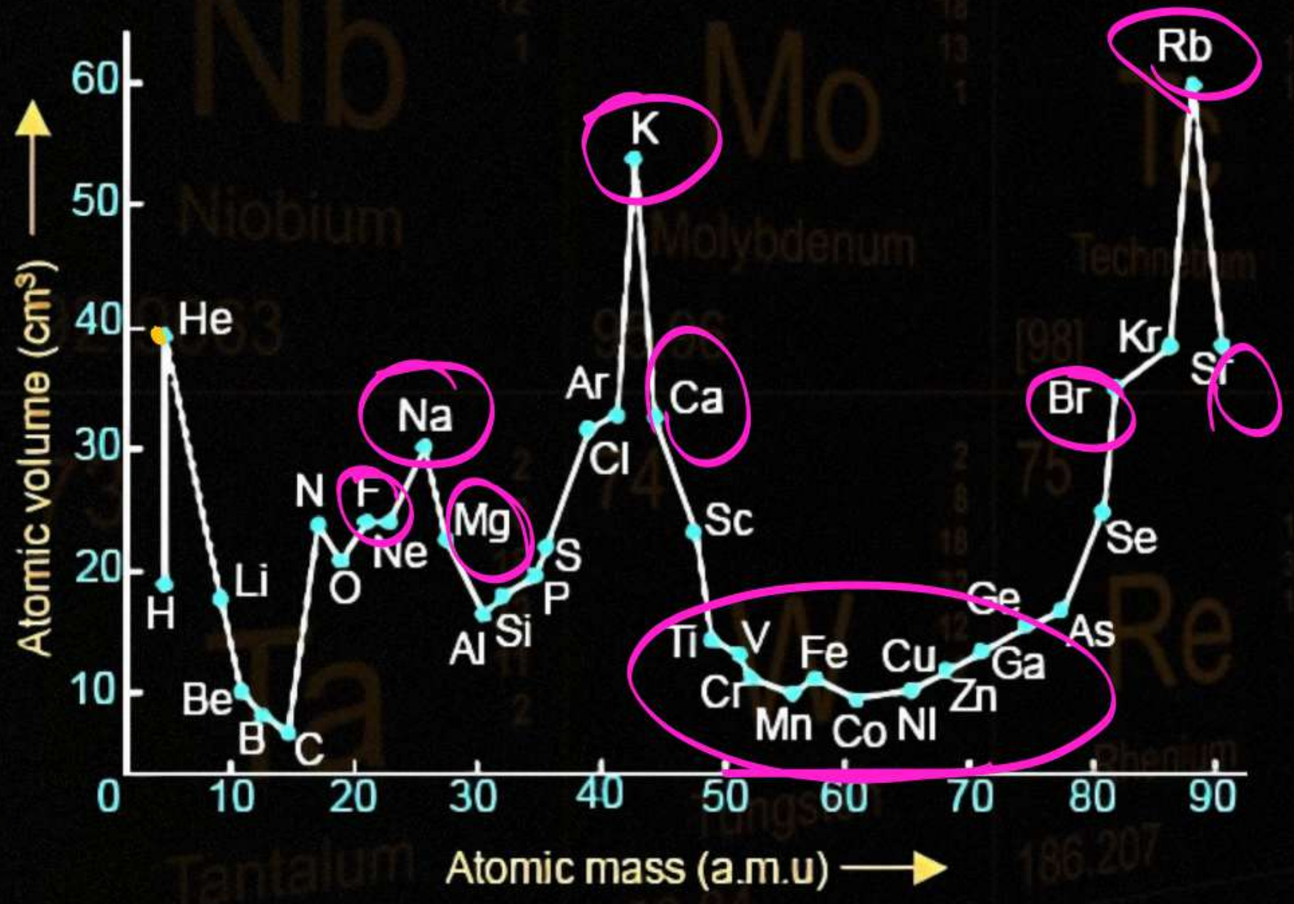
$\frac{7+39}{2} = \frac{46}{2} = 23$

| 1st | Li | Be | B | C | N | O | F |
|---|---|---|---|---|---|---|---|
| 8th | Na | Mg | Al | Si | P | S | Cl |

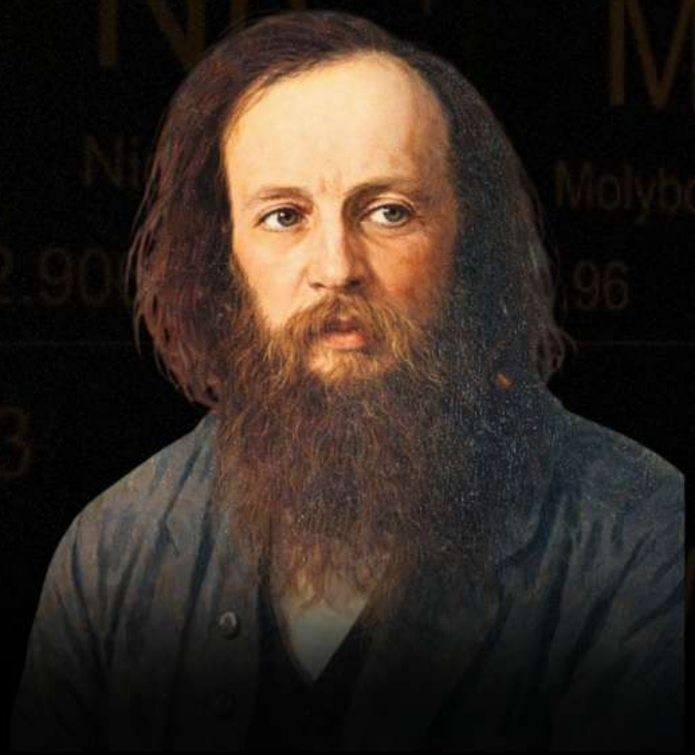


Top → Alkali Metal
Ascending Order → Halogen
Descending Order → Alkaline Earth Metals
Bottom → transition

# Lothar Meyer Atomic Volume Curve

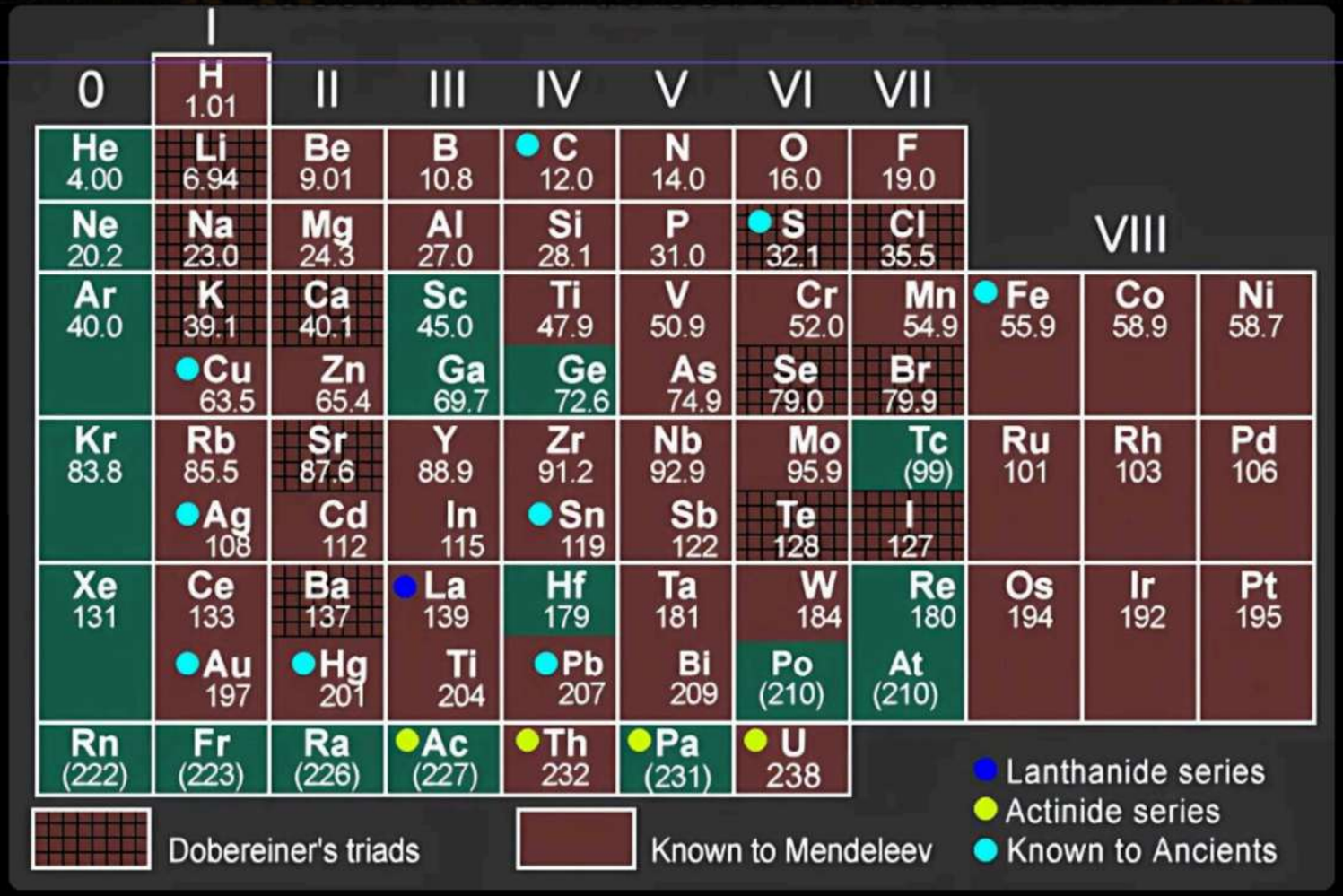


# Mendeleeve's Periodic Law



- It states that the physical and chemical properties of elements repeat periodically when arranged in order of increasing atomic weight.
- This means that elements with similar properties will appear after regular intervals of increasing atomic weight.
- Placed 63 elements in rows & columns.

# Mendeleeve's Periodic Law



# Merits of Mendeleev Periodic Table

- Some gaps were left for the elements yet to be discovered. Thus, if a certain new element is discovered, it can be placed in a new group without disturbing any existing group.

# Demerits of Mendeleev Periodic Table

- He was unable to locate hydrogen in the periodic table.
- Increase in atomic mass was not regular while moving from one element to another. Hence, the number of elements yet to be discovered was not predictable.
- Later on, isotopes of elements were found which violated Mendeleev's periodic law.

# Eka-aluminium and Eka-silicon

- In Dmitri Mendeleev's periodic table, the predicted elements below aluminum and silicon were called "eka-aluminum" and "eka-silicon," respectively.
- These elements were later discovered and named gallium and germanium, respectively.

| S.NO | Property | Eka-aluminium (Predicted) | Gallium (Observed) | Eka-silicon (Predicted) | Germanium (Observed) |
|---|---|---|---|---|---|
| 1. | **Atomic weight** | 6.8 | 70 | 72 | 72.59 |
| 2. | **Density ($g/cm^3$)** | 5.9 | 5.94 | 5.5 | 5.35 |
| 3. | **Melting point** | low | 29.78°C | High | 947°C |
| 4. | **Formula of oxide** | $E_2O_3$ | $Ga_2O_3$ | $EO_2$ | $GeO_2$ |
| 5. | **Formula of chloride** | $ECl_3$ | $GaCl_3$ | $ECl_4$ | $GeCl_4$ |

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# Modern Periodic Law



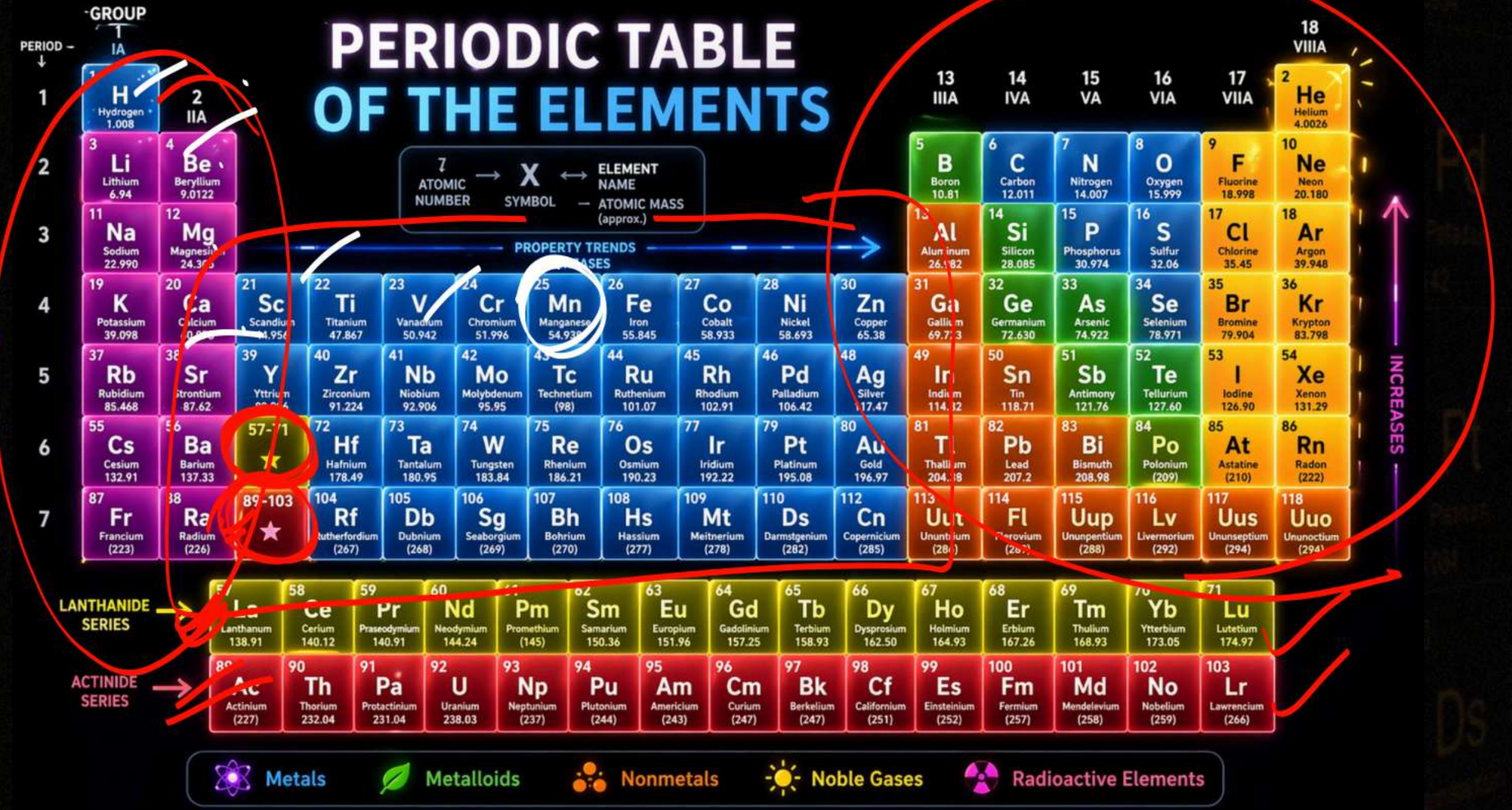
"The physical and chemical properties of the elements are periodic function of their atomic numbers."

Henry Moseley

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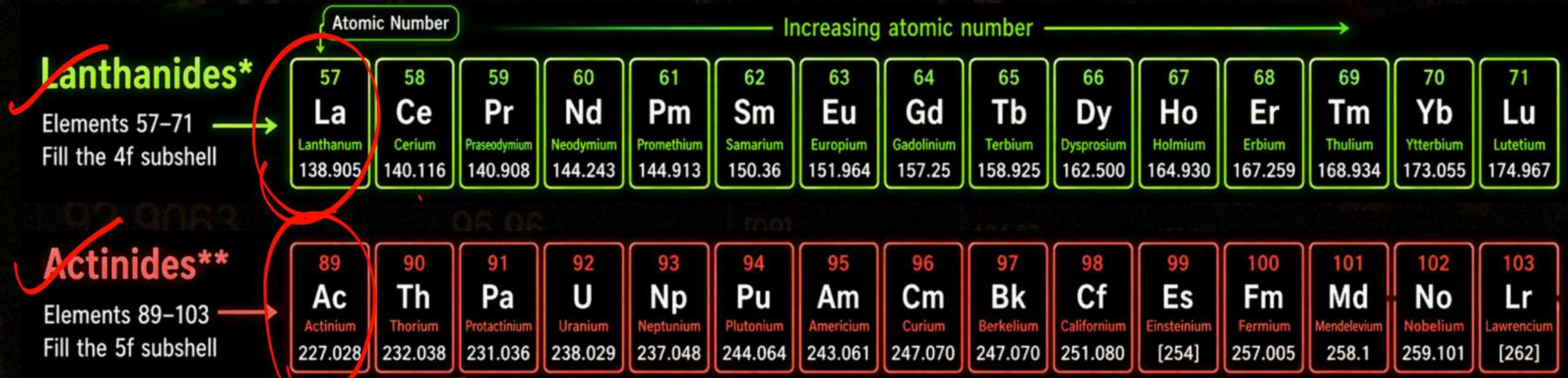
# Long form of periodic table



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# From where Lanthanum & Actinum series came ?



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# IUPAC Nomenclature of Elements with Z>100

| Digit | Name | Abbreviation |
|---|---|---|
| 0 | nil | n |
| 1 | un | u |
| 2 | bi | b |
| 3 | tri | t |
| 4 | quad | q |
| 5 | pent | P |
| 6 | hex | h |
| 7 | sept | S |
| 8 | oct | o |
| 9 | enn | e |

1 4 3 → uqt

un quad trium.

unquadtrium.

1 0 1

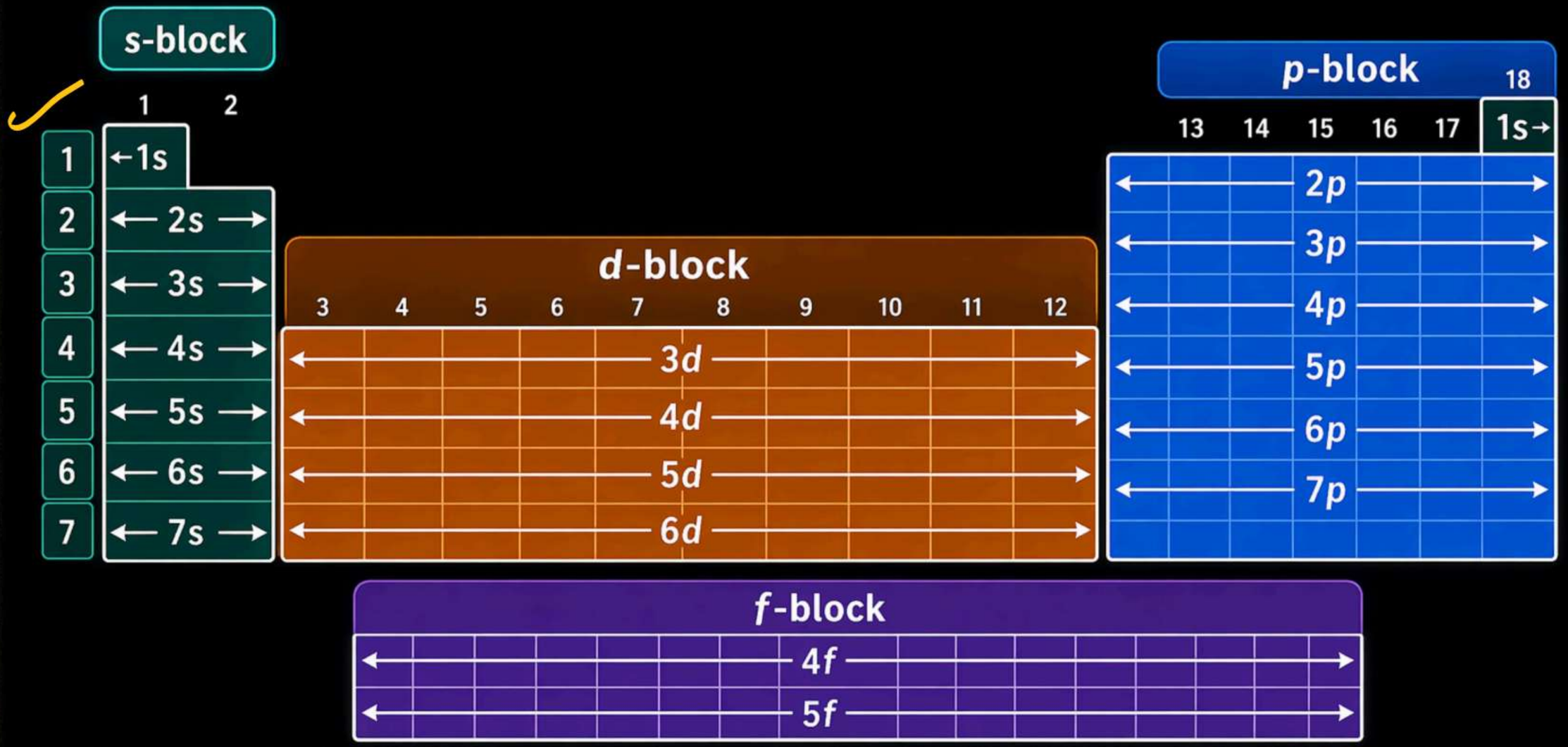
un nil un tium

# IUPAC Nomenclature of Elements

| Atomic Number | Name according to TUPAC nomenclature | Symbol | IUPAC Official Name | IUPAC Symbol |
|---|---|---|---|---|
| 101 | Unnilunium | Unu | Mendelevium | Md |
| 102 | Unnilbium | Unb | Nobelium | No |
| 103 | Unniltrium | Unt | Lawrencium | Lr |
| 104 | Unnilquadium | Unq | Rutherfordium | Rf |
| 105 | Unnilpentium | Unp | Dubnium | Db |
| 106 | Unnilhextum | Unh | Seaborgium | Sg |
| 107 | Unnilseptium | Uns | Bohrium | Bh |
| 108 | Unniloctium | Uno | Hassium | Hs |
| 109 | Unnilennium | Une | Meitnerium | Mt |
| 110 | Ununnillium | Uun | Darmstadtium | Ds |
| 111 | Unununnium | Uuu | Rontgenium | Rg |
| 112 | Ununbtum | Uub | Copernicium | Cn |
| 113 | Ununtrium | Uut | Nihonium | Nh |
| 114 | Ununquadium | Uuq | Flerovium | F1 |
| 115 | Ununpentium | Uup | Moscovium | Mc |
| 116 | Ununhextum | Uuh | Livermorium | Lv |
| 117 | Ununseptium | Uus | Tennessine | Ts |
| 118 | Ununoctium | Uuo | Oganesson | Og |

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# Electronic Configuration ( S,p,d & f )

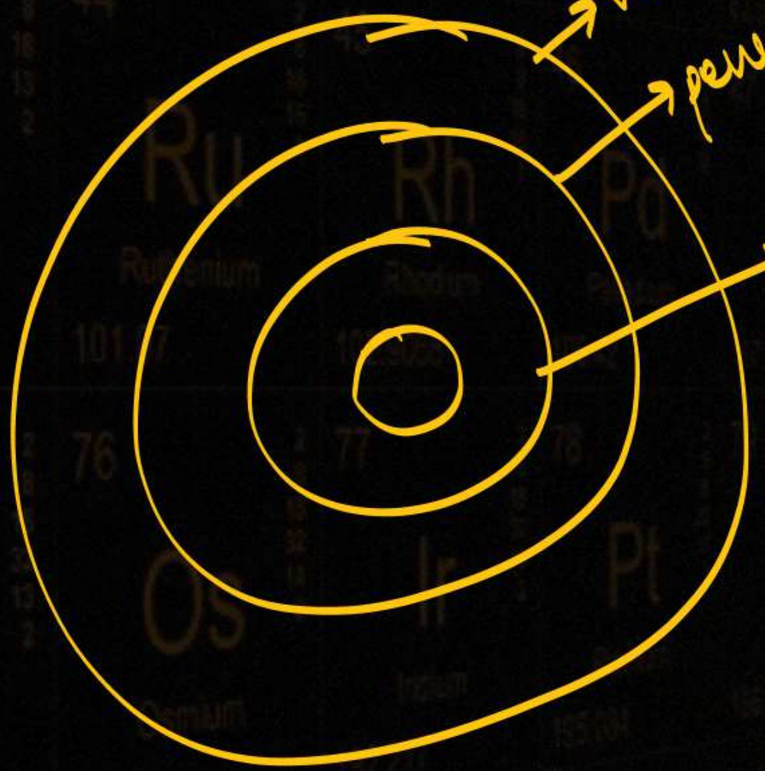
$Na \rightarrow 1s^2 2s^2 2p^6 3s^1$

| s-block elements | d-block elements | p-block elements |
|---|---|---|
| $ns^{1-2}$ | $(n-1)d^{1-10}ns^{1-2}$ | $ns^{1-2}np^{1-6}$ |

f-block elements

$(n-2)f^{1-14}(n-1)d^{0-1}ns^2$

$Sc \rightarrow [Ar]4s^2 3d^1$



# Bridge Elements

- Bridge elements are the elements of the second period of the periodic table.
- These elements show a relationship with the third-periodelements or the typical elements, because of similarity in their properties.
- They show similarities in characteristics such as atomic size, electronegativity and ion polarization.

| Group | 1 | 2 | 13 | 14 |
|---|---|---|---|---|
| Period 2 | Li | Be | B | C |
| Period 3 | Na | Mg | Al | Si |

# To Predict the Period, Group and Block of An Element From Its Electronic Configuration

Write the electronic configuration of the element.

1. **Period of the element :** It is equal to the principal quantum number (n) of the outermost orbit (shell).
2. **Block of the element**: It corresponds to the sub-shell in which last electron enters.



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# Group number of the element: It is determined by the following rules:

### (i) For s-block elements

Group number is equal to the number of electrons in the outermost s - orbital.

### (ii) For p-block elements

Group number = 12 + total number of electrons in outermost shell.

### (iii) For d-block elements

Group number = electrons in outermost (n) s - orbital + electrons in (n - 1)d - orbital.

## S block

$Na \rightarrow 1s^2 2s^2 2p^6 3s^1$

Period – 3rd

Block – s block

Group – I

## p-block

$Al \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^1$

Period – 3rd

Block – p-block

Group – 12 + ne⁻
12 + 1 ⇒ XIII th

## d-block

$Mn \rightarrow [Ar] 4s^2 3d^5$

Period → ~~3rd~~ → 4th

Block → d block

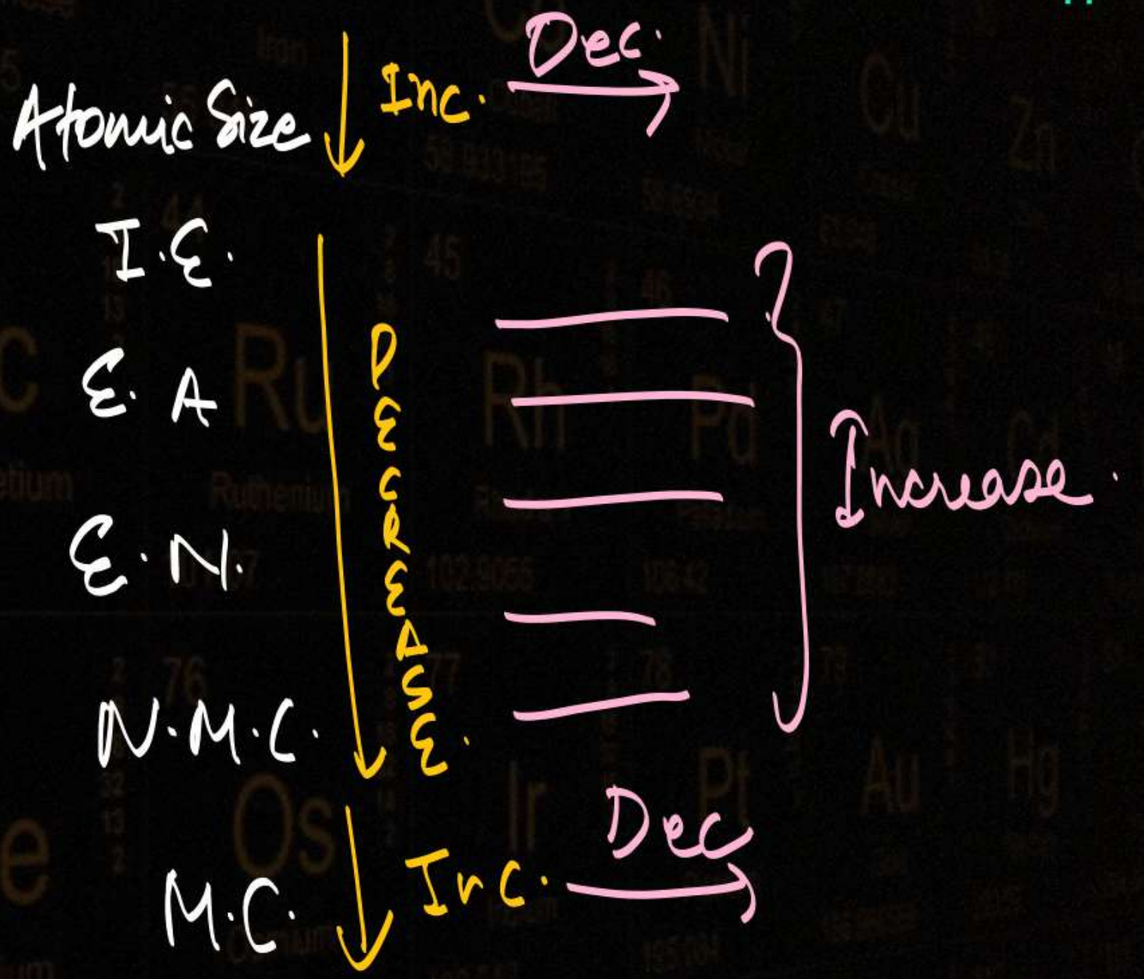
Group → 2 + 5 → 7th VII.

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# Periodic Properties

- Atomic Radius
- Ionisation Energy
- Electron Affinity
- Electronegativity
- Metallic Character
- Non Metallic Character



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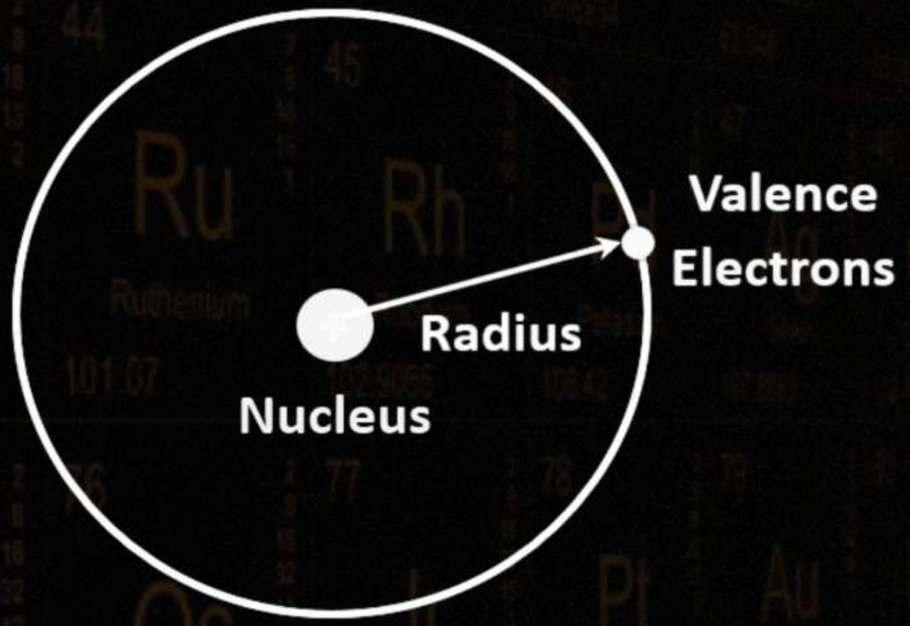
# IS IT EASY TO MEASURE

## Atomic Radius ?



# Atomic Radius

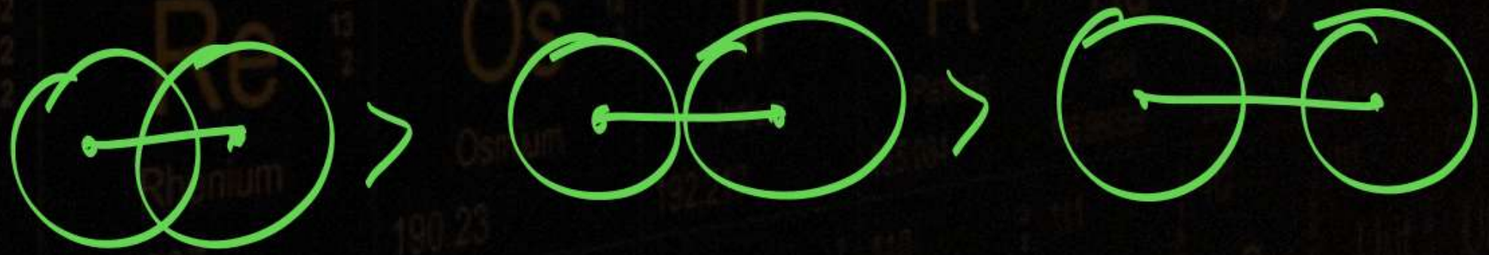
- Atomic Radii isthe total distance from the nucleus of an atom to the outermost orbital of its electron.
- The mean or typical distance from the centre of the nucleus to the boundary of the surrounding shells of electrons.



# Types of Atomic Radius with Respect to the Types of Bond

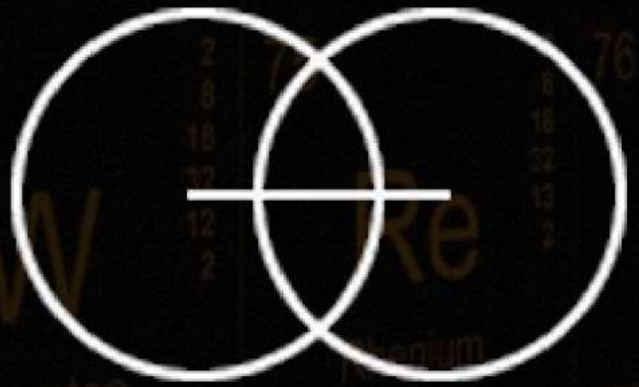
**Based on the type of bond, atomic radius is divided into three types as follows:**

- Covalent Radius →
- Vander Waals Radius →
- Ionic Radius → Anionic > Neutral > Cationic.
- Metallic Radius →



# Covalent Radius

- The covalent radius of an atom is the radius of an atom under the covalent bond with another atom(s) of a similar element.
- The covalent radius of an atom can be determined by measuring bond lengths between pairs of covalently-bonded atoms.
- If the two atoms are the same kind, the covalent radius is simply one-half of the bond length.



Covalent Radius

# Van Der Waals Radius

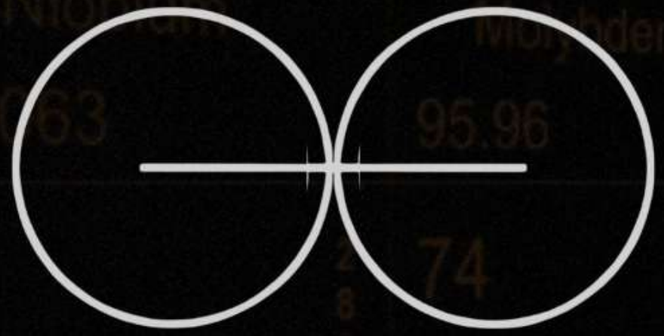
- Van-der-Waals radii are determined from the contact distances between unbonded atoms in touching molecules or atoms.



Vander Waal's Radius

# Metallic Radius

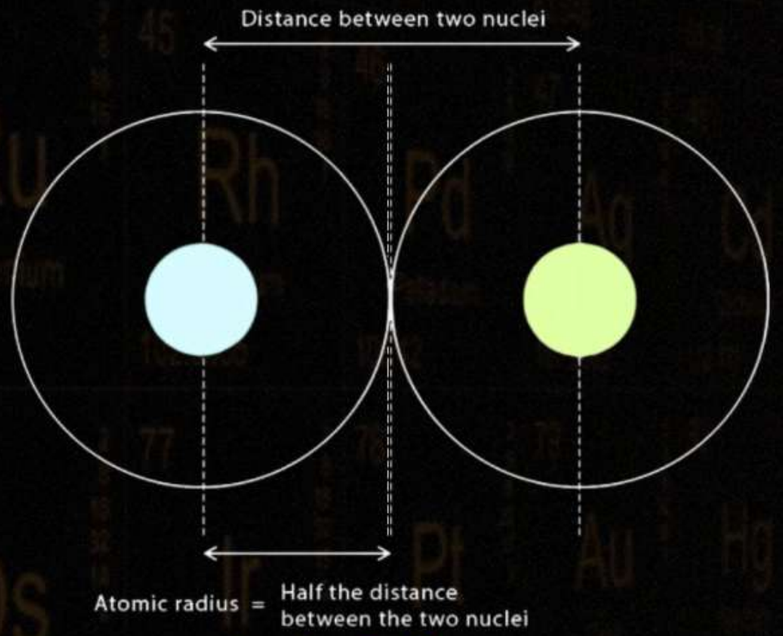
- The metallic radius is the radius of an atom joined by the metallic bond. The metallic radius is half of the total distance between the nuclei of two adjacent atoms in a metallic cluster



Metallic Radius

# Ionic Radius

- The ionic radius is the radius of an atom forming an ionic bond or an ion.
- The atomic bonds restrict the electrons and nucleus, and due to this, the ions or atoms don't have a specific shape.
- Anion > Neutral atom > Cation



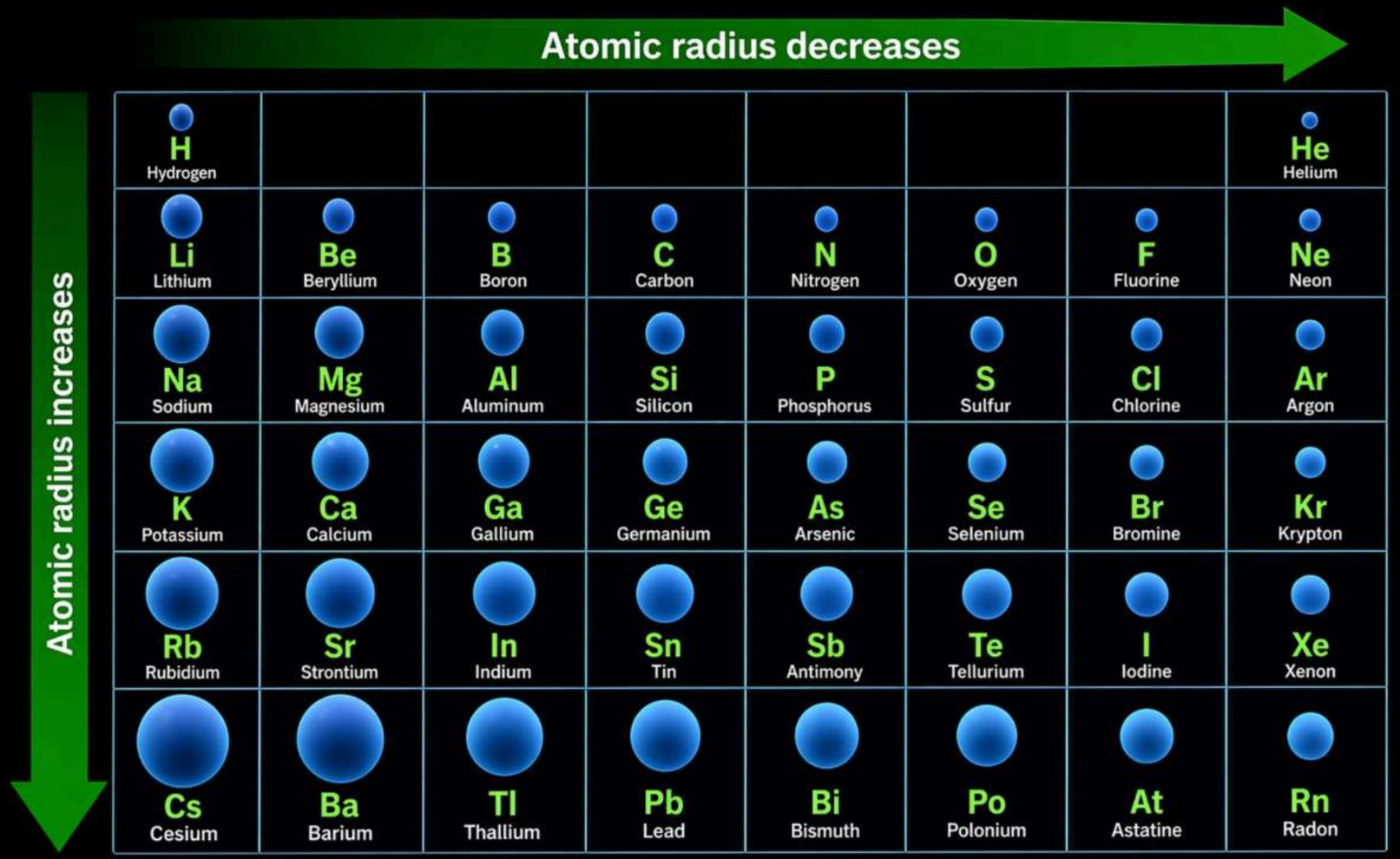
# Isoelectronic Specie

Isoelectronic species are atoms, ions, or molecules that have the same number of electrons. The term "isoelectronic" means "equal electric" or "equal charge".

## Example

- Ne- 10 electrons
- Na ion - 10 electrons
- F ion- 10 electrons

# Atomic Radius General Trend

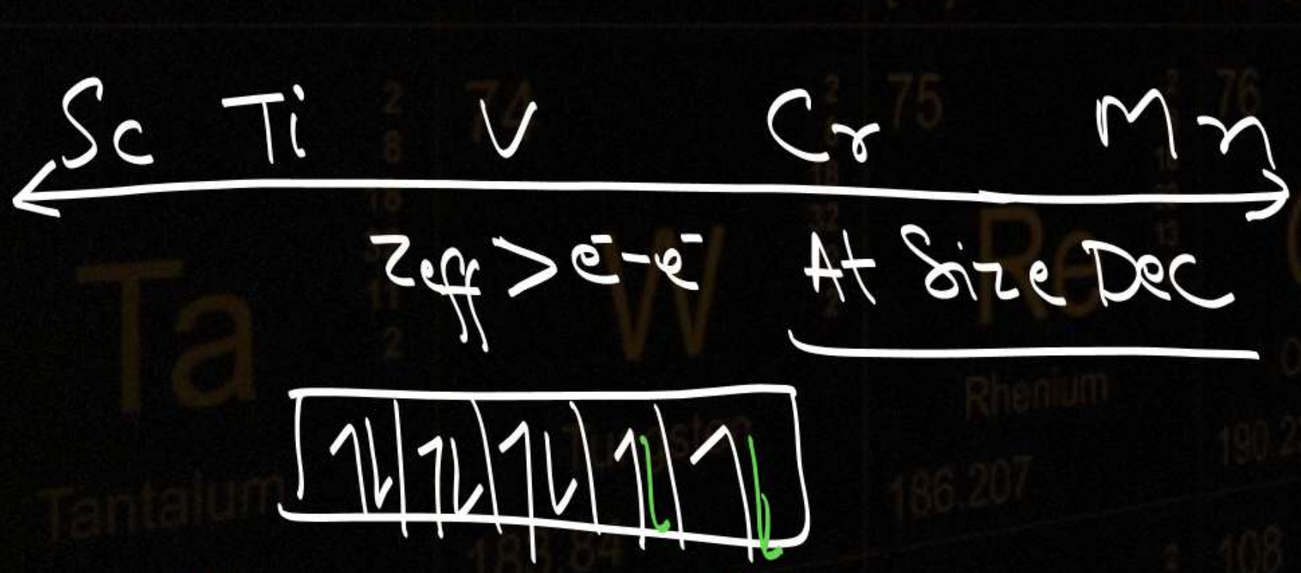


# Exceptions

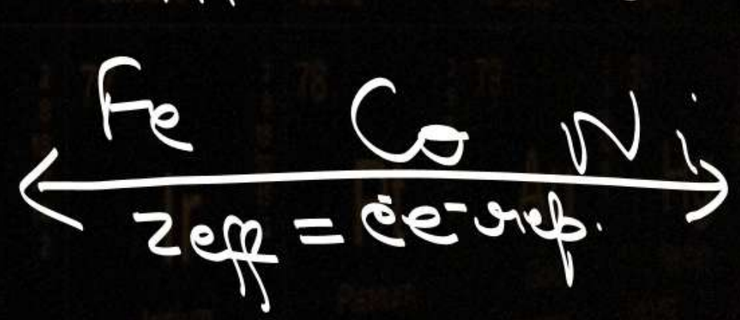
- In Period 1 : Atomic radius of He > H (Similar in all the periods).

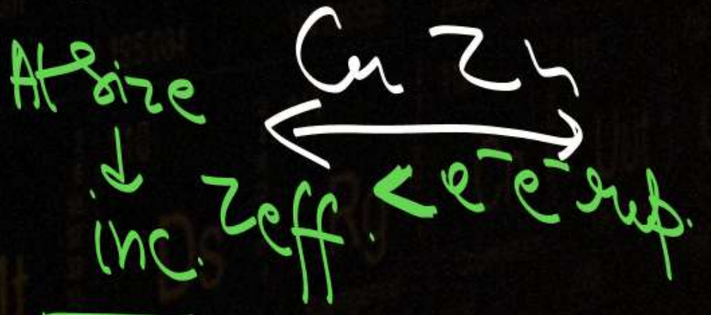
**Reason- Noble gases form Vander walls radius whereas the other elements form Covalent Radius.**

- D- Block elements show variable changes in atomic radius.



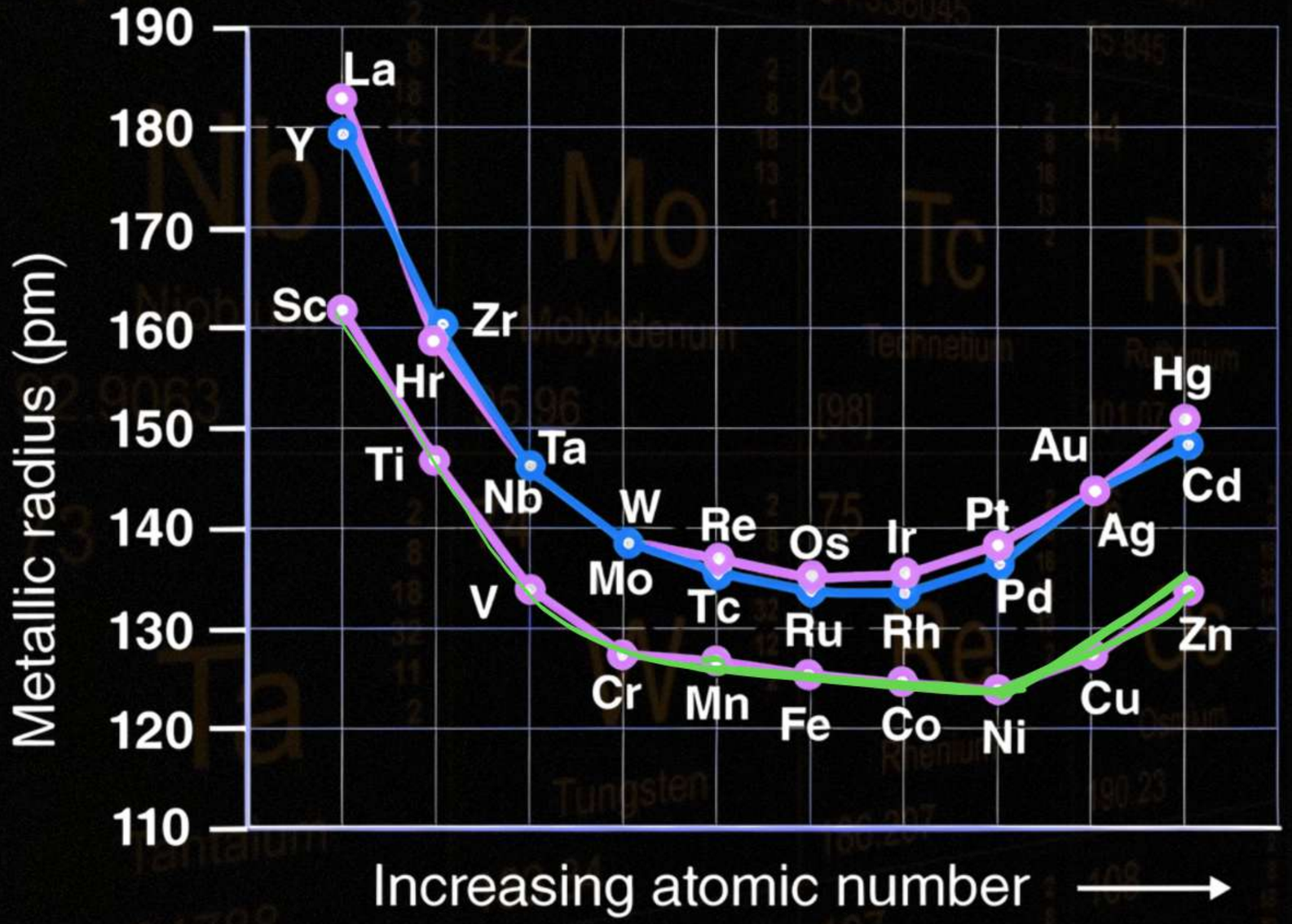


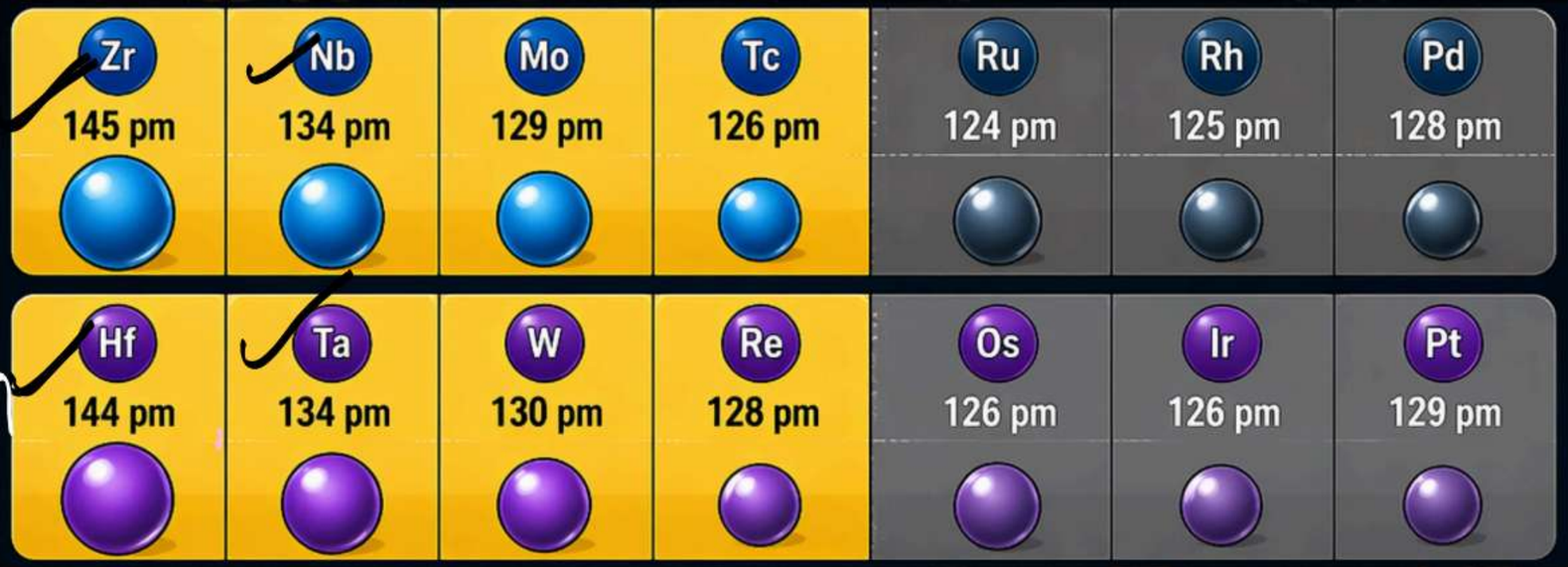



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# Down a Group

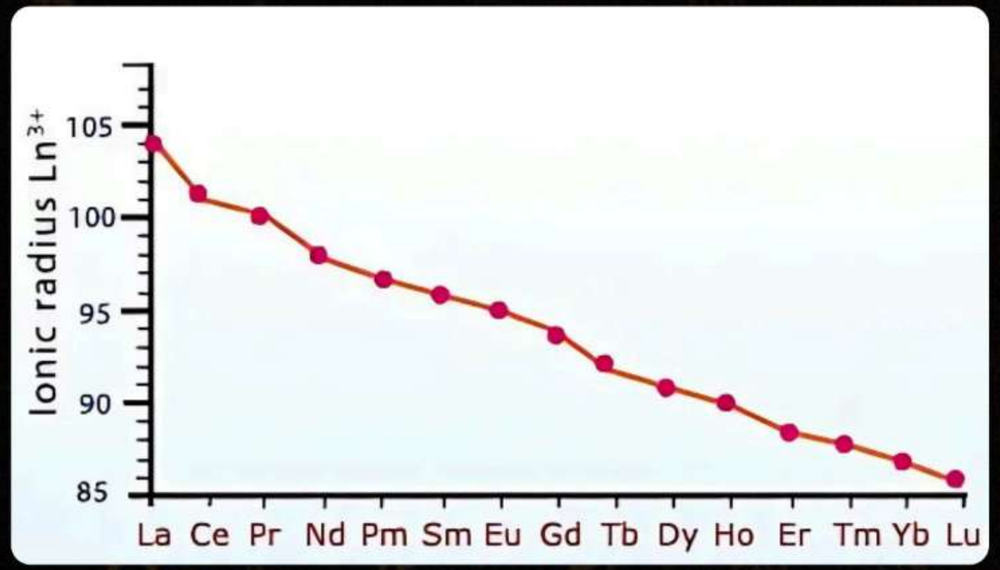



Ln

→ 14 Lanthanoids → poor S.E → Ln contr—

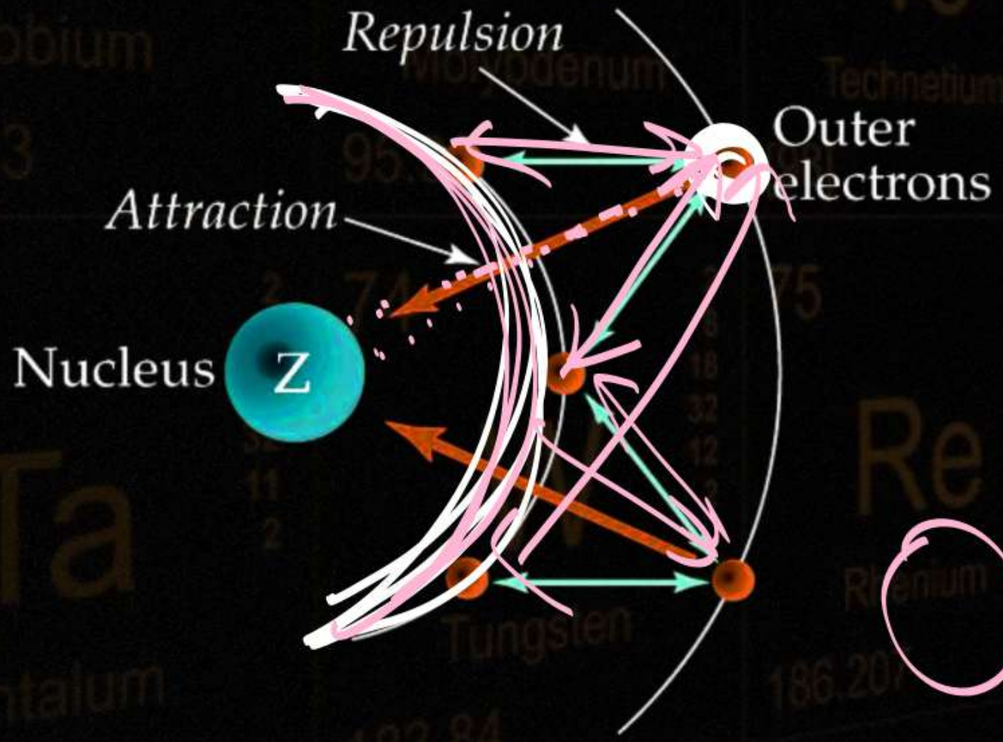
# Lanthanide Contraction

- Lanthanide contraction isthe steady decrease in atomic and ionic size of lanthanoids as their atomic number increases.
- Causes of Lanthanide contraction: In f-block elements, the positive charge on the nucleus rises by one unit, and one additional electron enters the same subshell, as the atomic number increases.



# Sheilding Effect

- The shielding effect can be defined asa reduction in the effective nuclear charge on the electron cloud, due to a difference in the attraction forces on the electrons in the atom.
- s>p>d>f




s>p>d>f

poor S.E.

# Ionization Energy

- The amount of energy required to remove an electron from an isolated atom or molecule. There is an ionization energy for each successive electron removed; the ionization energy associated with removal of the first (most loosely held) electron, however, is most commonly used.
- Ionization energy as the minimum energy that an electron in a gaseous atom or ion has to absorb to come out of the influence of the nucleus.
- It is also sometimes referred to as ionization potential and is usually endothermic process.
- It can also be used to determine the strength of chemical bonds. It is measured either in units of electronvolts/atom or kJ/mol.
- 1eV/atom = 96 kJ/mol

# Ionization Energy

- The first ionization energy is the energy that is required to remove the first electron from a neutral atom.
- The energy needed for the removal of the second electron away from the unipositive ion is second ionization energy, and so on.

## Example

$M + \Delta H_{1st} \rightarrow M+ + e;\ \Delta H1$ = First ionization energy

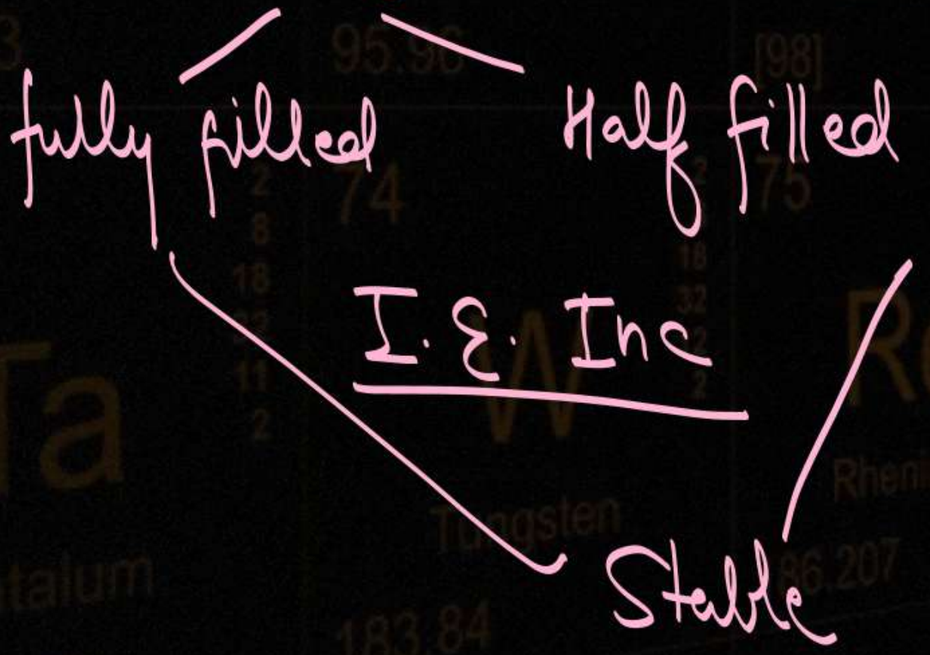
$M^{+} + \Delta H_{2nd} \rightarrow M_2+ + e;\ {}^{-}\Delta H_{2nd}$ = Second Ionization energy, etc.

# Ionization Energy

- Naturally, removing the second electron from an already positive ion will be difficult. Hence, the second ionization energy will be larger than the first ionization energy.
- The third ionization energy will be more than the second ionization energy etc.
- $\Delta H_{1st} < \Delta H_{2nd} < \Delta H_{3rd} < .....$ Because of the enhanced stability of half-filled and fully-filled .
- Orbitals, the removal of electrons from such systems will have relatively higher ionization than other atoms and ions.
- Highest Ionisation Energy - Helium
- Caesium element has the lowest ionisation energy.

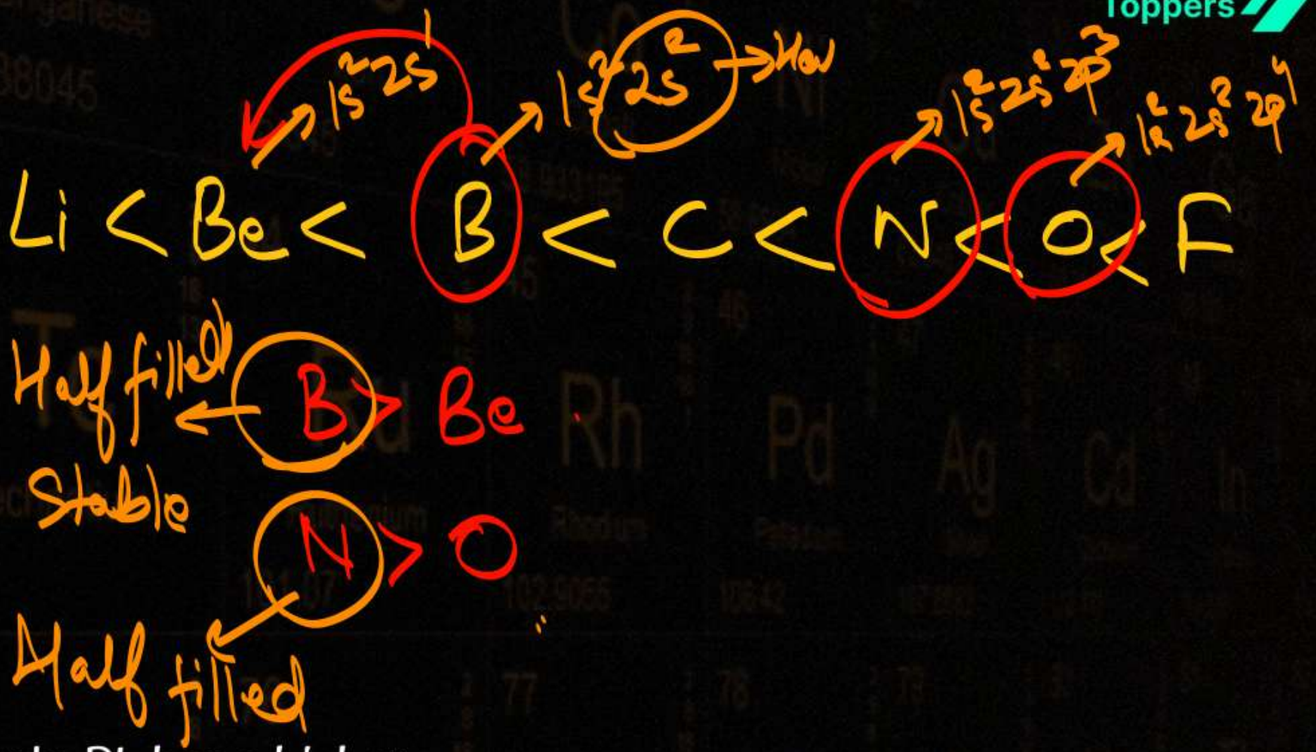
# Factors on which I.E. depends :-

- Atomic Size $\propto \frac{1}{I.E.}$
- Nuclear Charge $\propto I.E.$
- Electronic Configuration



# Exceptions

- **I.E. of Boron < Berilium**
- **I.E. of Nitrogen > Oxygen**
- **I.E. of Magnesium > Aluminium**
- **I.E. of Phosphorus > Sulphur**



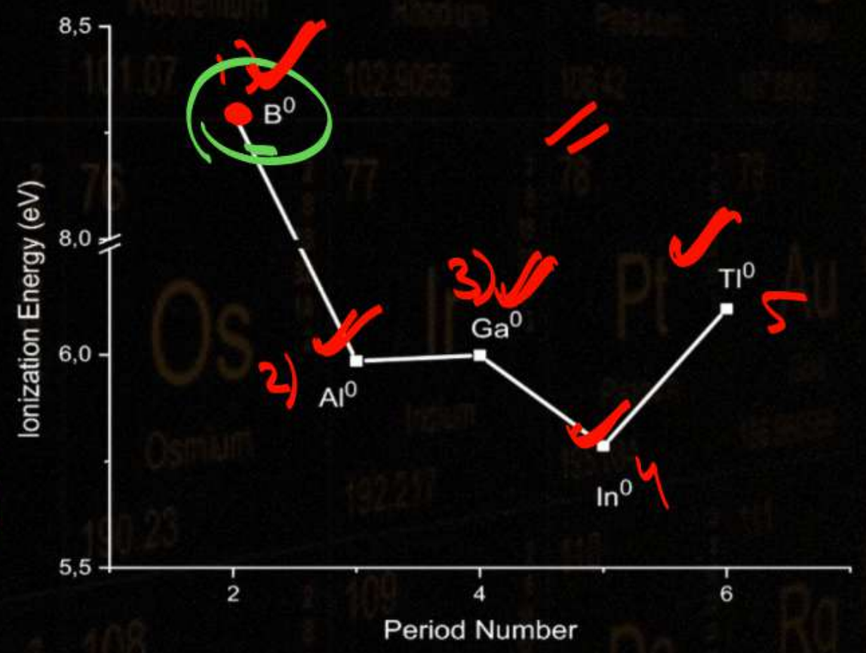

- C > Si > Ge > Pb >Sn.
- Due to ineffective shielding of the d and f orbitals, Pb has a higher Ionization Enthalpy than Sn in this case.

# Ionisation Energy Trend Exceptions in Group

**As we go from indium to thallium,screening by the filled 5d and 4f subshells is so poor that Zeff increases significantly from indium to thallium. Thus the first ionization energy of thallium is actually greater than that of indium.**

- B> Tl> Ga> Al> In
- Similarly in group 14, the I.E. of Pb is more than Sn
- C>Si>Ge>Pb>Sn



# Electron Affinity

- The electron affinity is the potential energy change of the atom when an electron is added to a neutral gaseous atomto form a negative ion.
- So the more negative the electron affinity the more favourable the electron addition process is.
- Not all elements form stable negative ions in which case the electron affinity is zero or even positive.
- It is an Exothermic Process
- It can also be used to determine the strength of chemical bonds. It is measured either in units of electronvolts or kJ/mol.

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# Successive Electron Affinity

EA (first) + EA ( second) < 0

# Electron Affinity

- The first electron affinity is generally exothermic that is negative.
- The second electron affinity of the same element will be positive or endothermic.
- This is so because the second electron has to be forced to enter the mono negative ion.
- The electron affinity cannot be determined directly but is obtained indirectly from the Born-Haber cycle.

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# Factors Affecting Electron Affinity

- 1. Atomic Size $\propto \frac{1}{E.A.}$
- 2. Nuclear Charge $\propto N.C.$
- 3. Electronic Configuration

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# Electron Affinity General Trend

# Exceptions

- Period-2 Ne < N < Be < B < Li < C < O < F.
- Group-17 Astatine< Iodine < Bromine < Fluorine < Chlorine
- Chlorine has highest Electron Affinity
- Fluorine is the most electronegative element but still its electron affinity is less than that of chlorine .
- It is because of the small size of fluorine. All the electrons in the atom repels the incoming electron significantly, this nullify the effect of higher electronegativity of fluorine over chlorine.
- Oxygen has the lowest electron affinity value in group 16
- Similar trend will be followed by the elements of group 14,15 & 16 i.e. E.A. of Si > C, P >N & S > O.
- Group 17 has always more electron affinity as compared to group 16 elements.

# Electronegativity

- It is the tendency of an atom in a molecule to attract the shared pair of electrons towards itselfis known as electronegativity.
- It is a dimensionless property because it is only a tendency.
- It basically indicates the net result of the tendencies of atoms in different elements to attract the bond-forming electron pairs. We measure electronegativity on several scales.
- The most commonly used scale was designed by Linus Pauling.

# Difference between Electronegativity & Electron Affinity

| Properties | Electronegativity | Electron Affinity |
|---|---|---|
| Definition | Tendency of an atom to attract electrons | Amount of energy released when an atom accepts an electron |
| Defined for | Covalent Bond | Isolated Atom |
| Units | Pauling | eV or kJ/mol |
| Value | Gives relative value | Gives absolute value |
| Higest value | Fluorine | Chlorine |

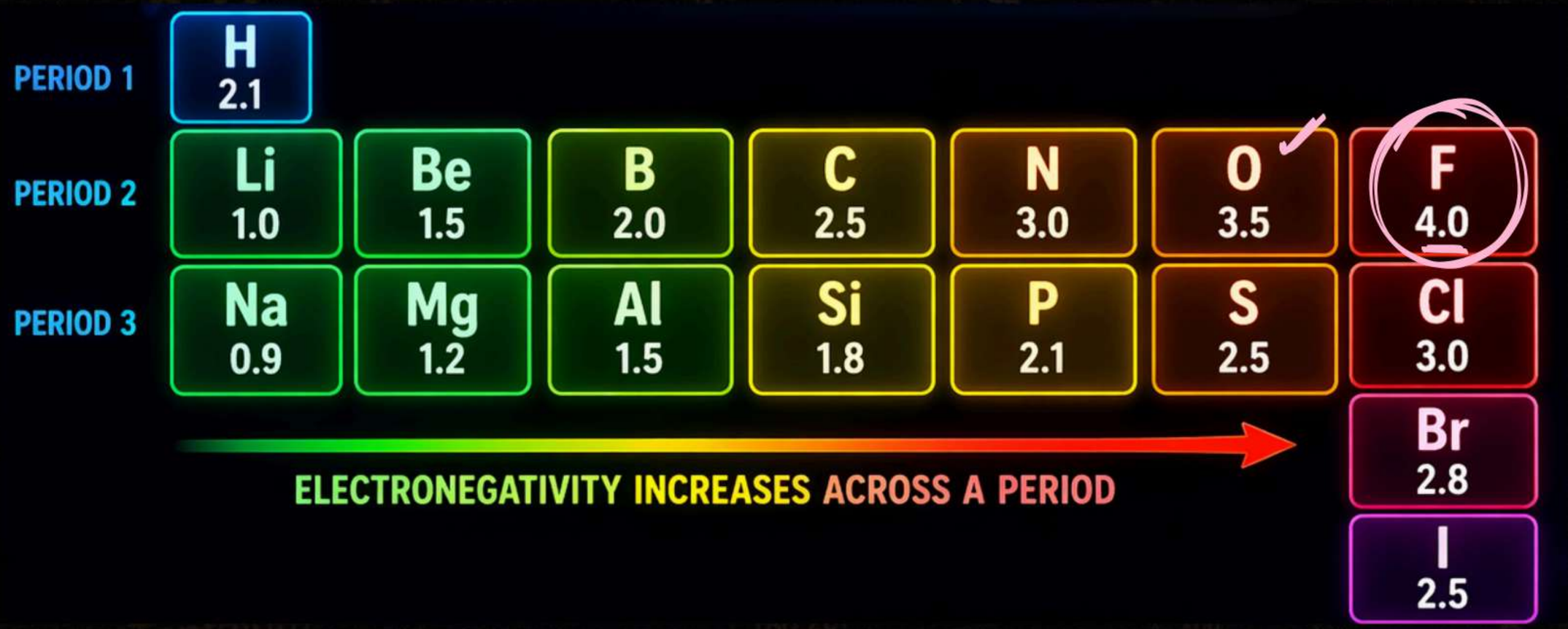
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# Factors Affecting Electronegativity

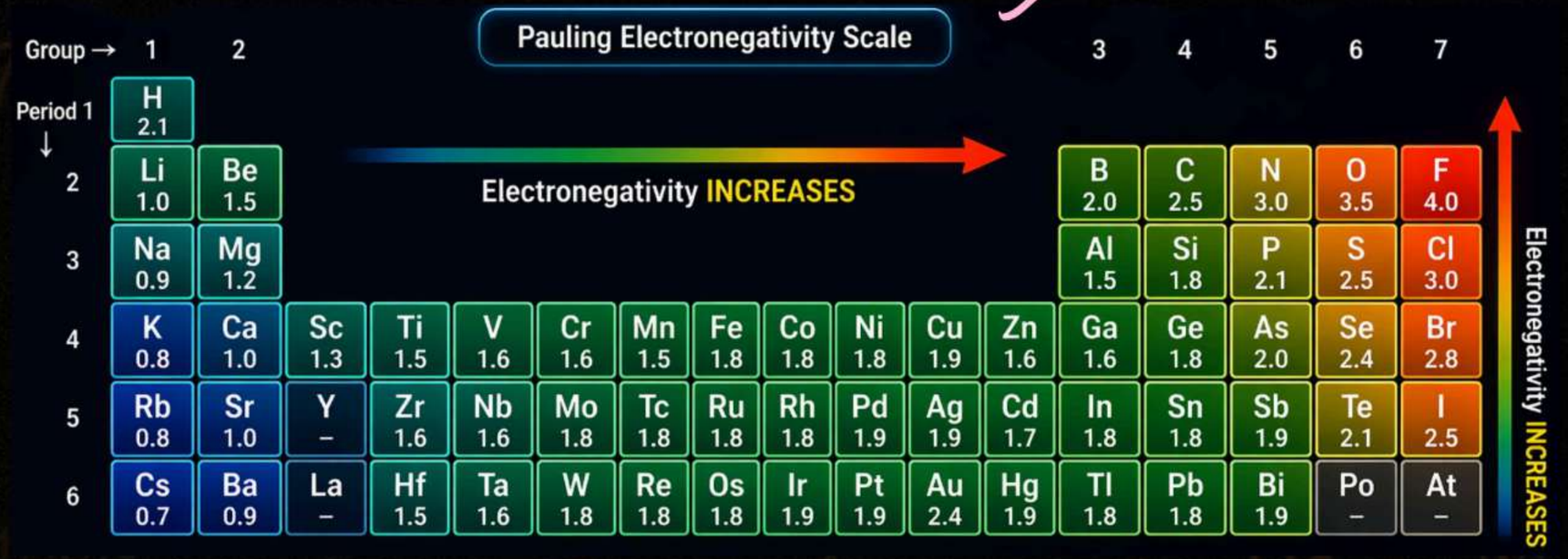
1. Atomic Size $\propto \frac{1}{E.N}$
2. Nuclear Charge $\propto E.N$
3. Charge on cation & anion
4. Hybridisation

# Electronegativity Values



# Electronegativity

- Fluorine is the most electronegative element with a value of 4.0.
- Cesium is the least electronegative element with a value of 0.7
- It is a general observation that metals show a lower value of electronegativity as compared to the non-metals.Therefore, metals are electropositive and non-metals are electronegative in nature.
- Electronegativity has no units. It is measured in scale.




Apne
itne acche se
pdha
23:05
restart
isliye meri taraf se
5 minute ka
break lelo
Keep Going!
Proud of you
You Did Great!
Break Time
Hard Work Today
Better Tomorrow
Future is Bright
Notes

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# Class 11th | Chemistry

## Chemical Bonding & Molecular Structure

## Chapter 4

7 Marks

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# Different Bonds

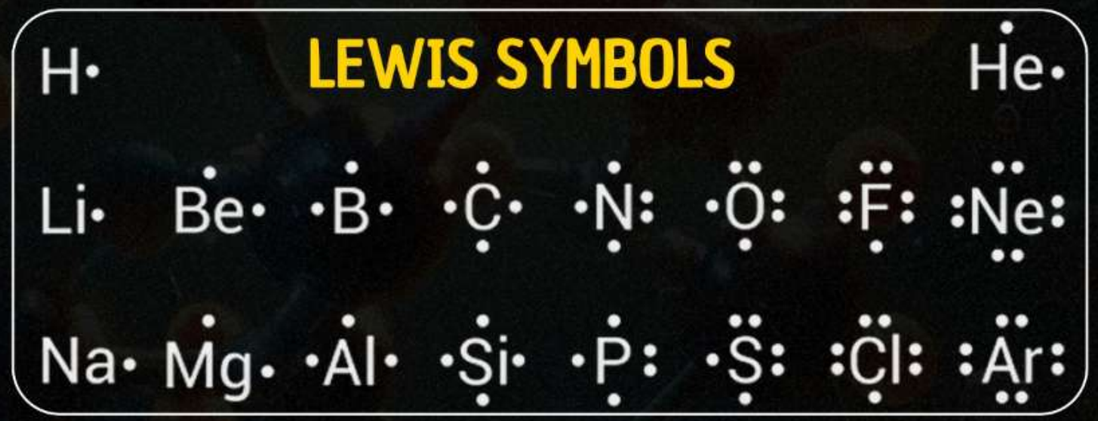


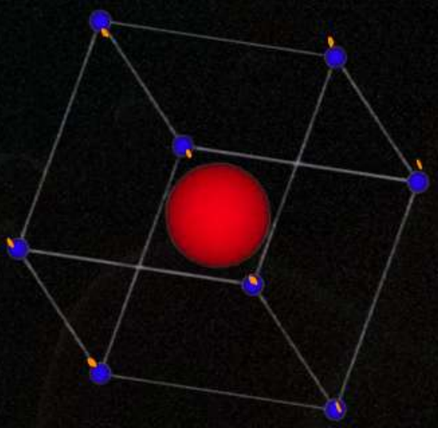
# Octet rule




- Atoms of all the atoms combine to complete their octet or duplet. It is called Octet Rule.
- Kossel-Lewis theory is based on the electronic concept of atom and is called electronic theory of valence.
- Kernel is +vely charged. The units of positive charge on kernel is equal to the number of valence electrons.

| Inert gas | Electronic configuration |
|---|---|
| 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 |



# Covalent bond

A covalent bond is formed by the mutual sharing of electron pairs among the combining atoms.



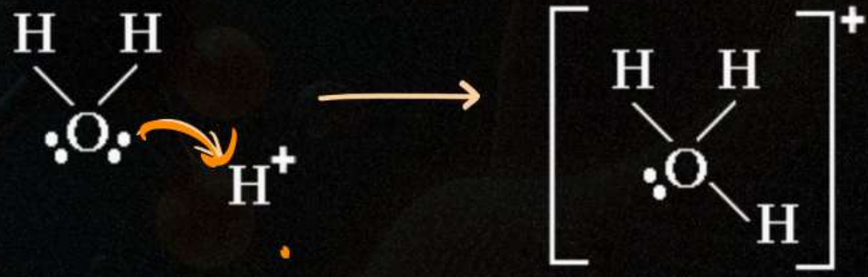

# Electrovalent Bond

It is the bond formed by the transfer of electron between a metal & a non metal .




# Coordinate bond or dative bond

It is a special type of covalent bond in which the shared electron pair is contributed by one of the combining atom only.



# Lewis dot structure (Octet Rule)

- Select the central atom which is generally the less electronegative atom or highest atomic number (sometimes) or less in numbers.
- Central atom should have atleast 8 electrons.
- Draw the skeleton & join them using single bond.
- Add the total number of valence electrons of each atom of the molecule.
- Put one pair of shared electrons between the atoms linked by a single bond, two pairs of shared electrons between the atoms linked by a double bond and so on.
- The remaining electrons are shown as lone pairs.
- The octet of each atom (duplet for H, Be) should be complete.

**Total no. of electrons (Q) = No. of valence e + ( -ve ) – (+ve ) charge.**

**Bond pair electron = 2 x no. of bonds.**

**Lone pair electron = Q - B.p e.**

**Lone pair electron is either used to fill lone pair or as double , triple bond according to the compound.**

# Covalency

- Covalency is the number of electron pairs an atom shares with other atoms to achieve a stable electronic configuration. It represents the total number of covalent bonds & lone pair electrons (if present) an element forms in a given molecule.



- Oxygen - 1,2
- Sulphur - 2,4,6
- Nitrogen - 3,4
- Phosphorous - 3,5
- Chlorine - 1,3,5,7

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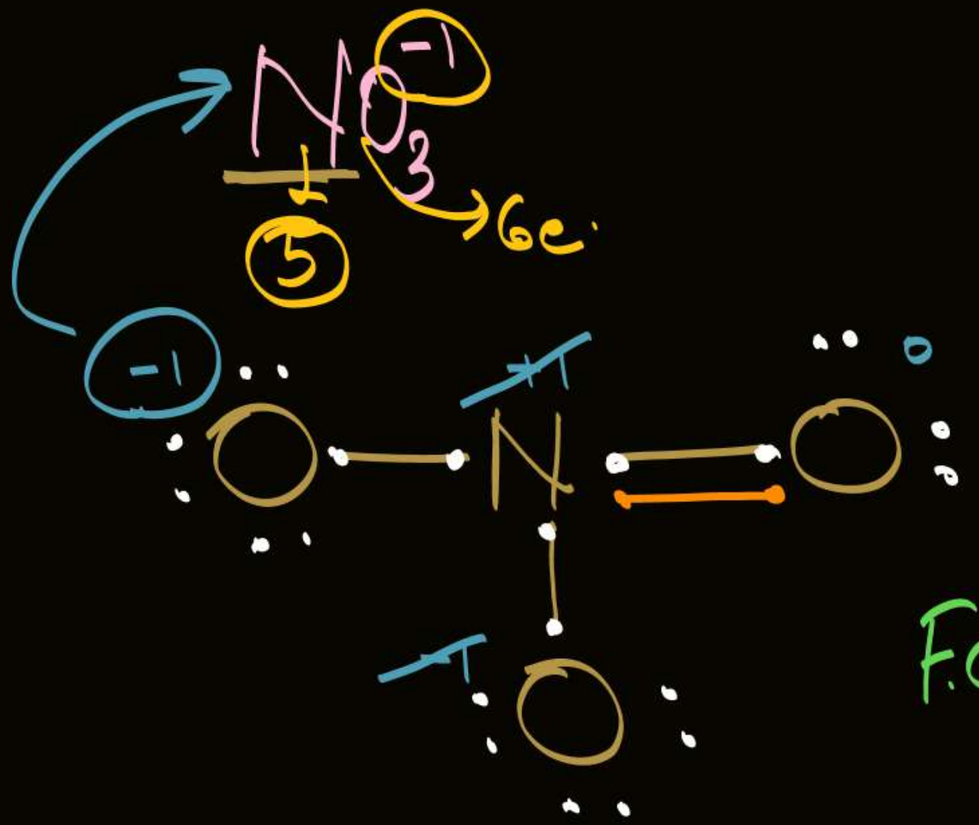
# Formal Charge

- Formal Charge is the difference between the number of valence electrons of that atom in free state (isolated state) and the number of electrons assigned to that atom in the Lewis structure.
- Signiticance of Formal Charge- Structure having smallest formal charge on the atoms is most stable. Thus we can select the most stable structure.

$$F.C = V - L - \frac{B}{2}$$

## Limitations of the Octet Rule

- Incomplete octet
- Odd-electron molecules
- Expanded octet
- Shape of the molecule
- Relative stability
- Hydrogen and helium
- Ionic and covalent bonds




$Q = 5 + 6 \times 3 + 1 \Rightarrow 24$

$B.p.e^- = 2 \times \text{Bonds} \Rightarrow 2 \times 3 \Rightarrow 6$

$L.p.e^- = Q - B$

$\Rightarrow 24 - 6 \Rightarrow 18$

$L.p. \Rightarrow 3, 3$

$F.C_{Nitrogen} = V - L - \frac{B}{2}$

$= 5 - 0 - \frac{8}{2} = +1$

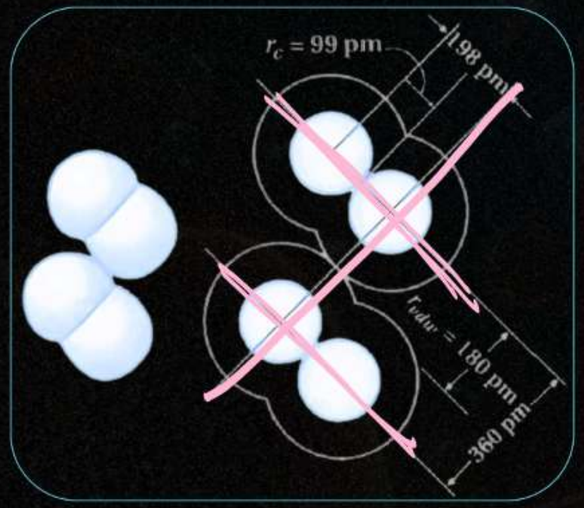
$F.C_{(-O)} = 6 - 6 - \frac{2}{2} \Rightarrow -1$

$F.C_{(=O)} \rightarrow 6 - 4 - \frac{4}{2} \Rightarrow 0.$

# Bond Parameters

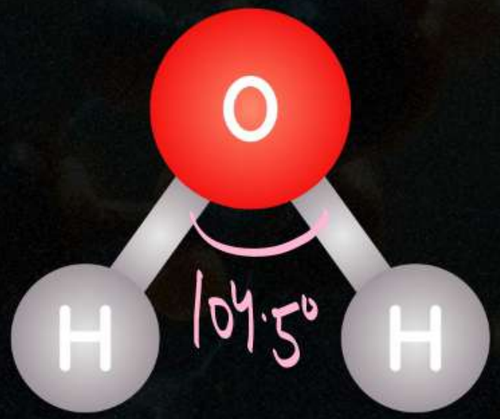
## 1. Bond Length

Bond length is defined as the equilibrium distance between the nuclei of two bonded atoms in a molecule.



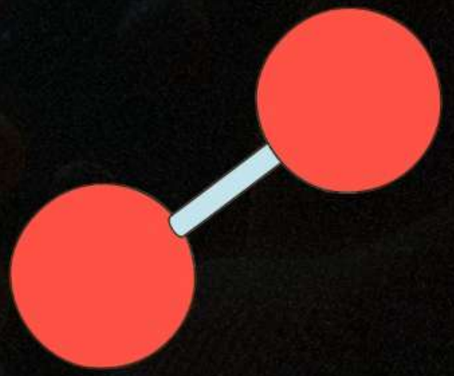

## 2. Bond Angle

It is defined as the angle between the orbitals containing bonding electron pairs around the Central atom in a molecule.




## 3. Bond Order

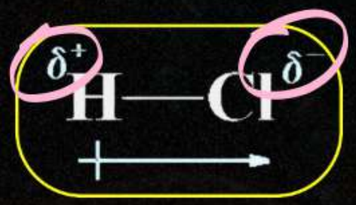
It is given by the number of bond between the two atoms in a molecule. The bond order is given by the number of bonds between the two atoms in a molecule.



# Bond Parameters

## 4. Bond Polarity

A dipole moment is a measurement of the separation of positive and negative electrical charges within a system.
It is a vector quantity that has both magnitude and direction



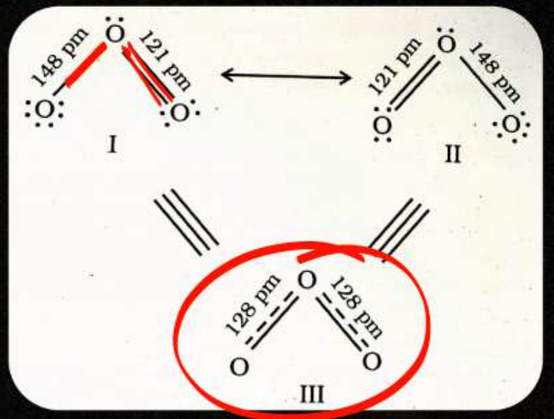

D.M. is Expressedin Debye

1 D = 3.33564 X $10^{-30}$ Cm.

**D.M. = charge(Q) X distance of separation(r)**

## 5. Resonance Structures

A single Lewis structure cannot describe a molecule accurately.
A number of structures with similar energy positions of nuclei, bonding and non-bonding, pairs of electron are taken as the canonical structure of the hybrid, which describes the molecule accurately.Resonance hybrid has less energy hence they are more stable.

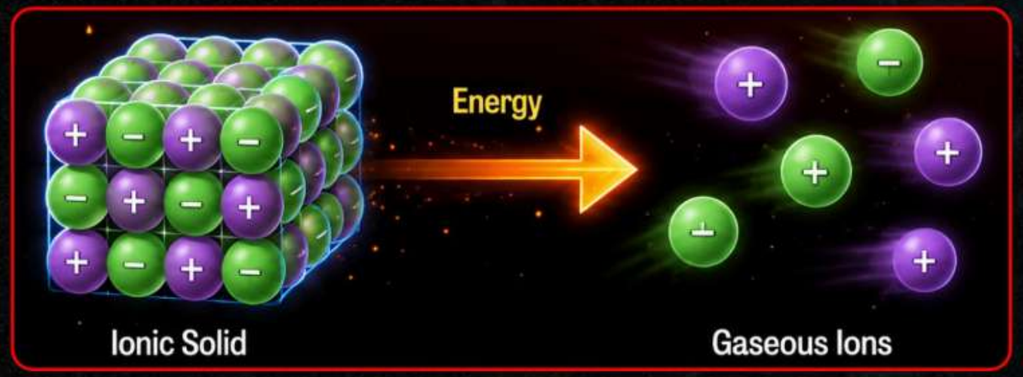



Resonance in ozone

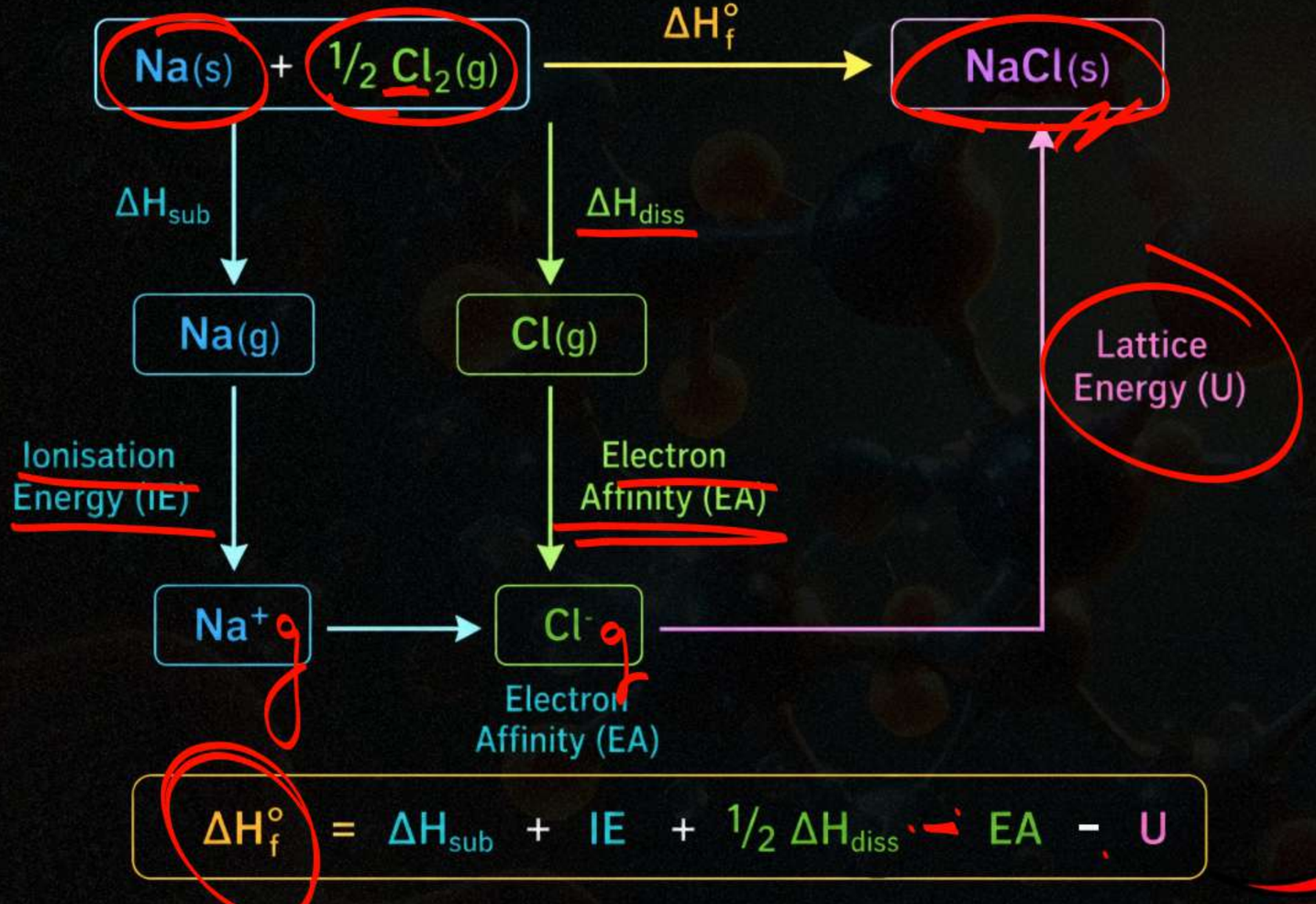


# Lattice Enthalpy

- **Lattice Dissociation Enthalpy**: The energy required to completely separate one mole of a solid ionic compound into its constituent gaseous ions. This process is **endothermic** (energy is absorbed, $\Delta H > 0$).
- **Lattice Formation Enthalpy**: The energy released when one mole of a solid ionic compound is formed from its constituent gaseous ions. This process is **exothermic** (energy is released, $\Delta H < 0$).
- **Born Haeber Cycle** is an indirect method to calculate lattice energy.



# Lattice Enthalpy




## Hess Law

# Fazan's Rule

All the covalent bonds have some partial ionic character and partial covalent character .When a cation approaches an anion, the electron cloud of the anion is attracted towards the cation and hence gets distorted. This is called Polarisation of the anion.

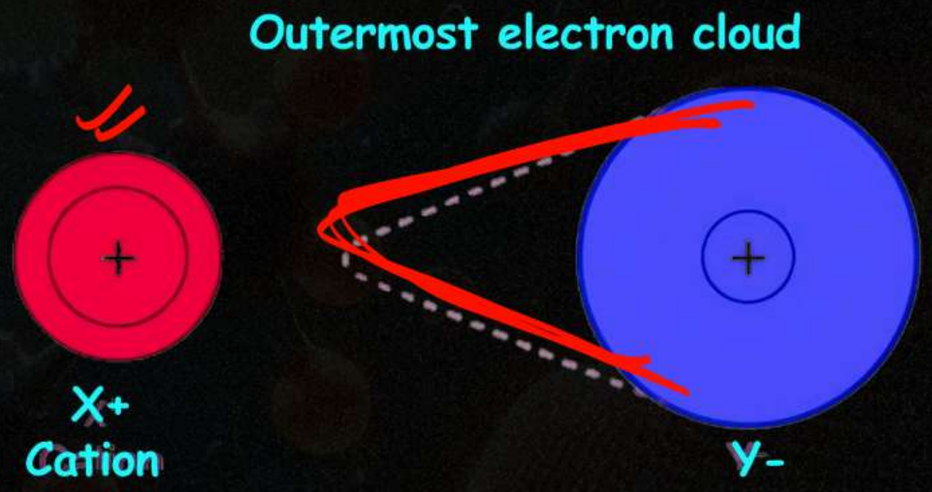
## Polarising power

Power of cation to polarise anion

- Smaller the cation, greater is its polarising power
- Larger the anion, greater is its polarisability
- Larger the charge on cation, more is its polarising power
- Larger the charge on anion, more easily it gets polarised.
- Same size & charge cations, then one with EC of transition metals is more polarising than one with noble gas, alkali or alkaline earth metal cations.

## Polarisability

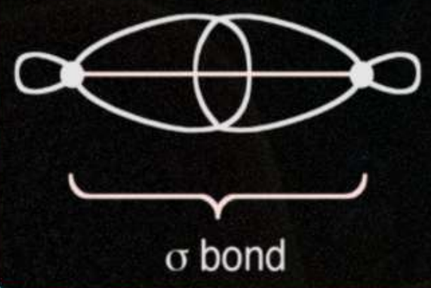
Tendency of anion to get polarised



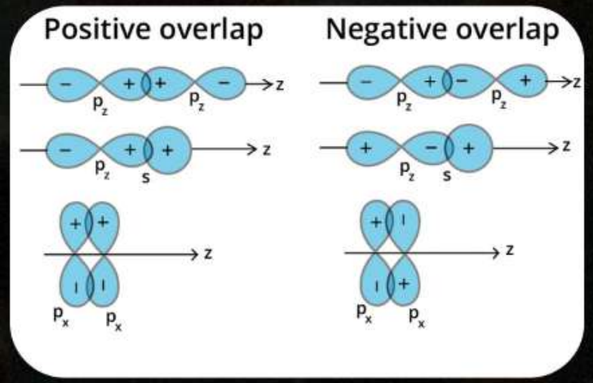
# Valence Bond Theory

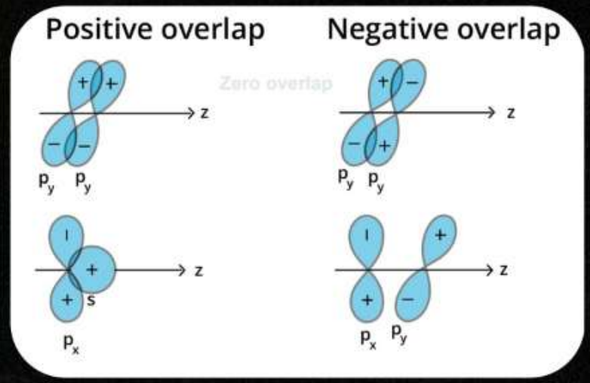
1. The theory assumes that electrons occupy atomic orbital's of individual atoms within a molecule, and that the electrons of one atom are attracted to the nucleus of another atom.
2. Greater the extent of overlapping, stronger is the bond formed.
3. The direction of the covalent bond is along the region of overlapping of the atomic orbital's i.e., covalent bond is directional.

| σ-bond | л-bond |
|---|---|
| A sigma bond (symbol: σ) is a covalent bond formed via linear overlap of two orbital's. | A pi bond (symbol: л) is a covalent bond formed via parallel overlap of two orbital's. |






ALWAYS
SIGMA

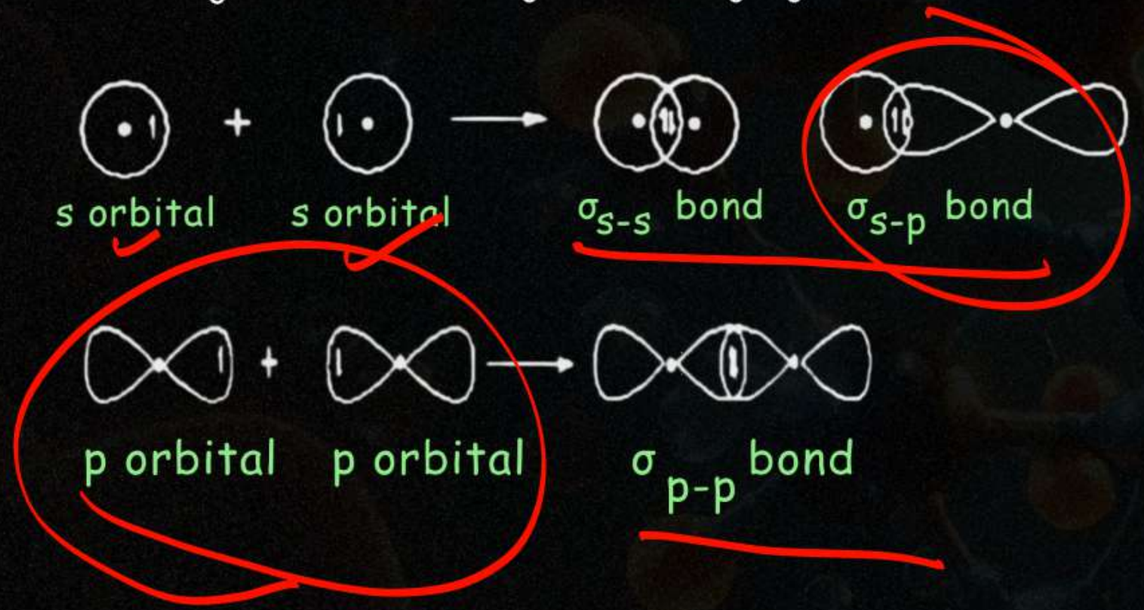
BACKLOG MARATHON SERIES

Next Toppers

# Valence Bond Theory

## σ-bond

The covalent bond formed due to overlapping of atomic orbital along the inter nucleus axis is called σ-bond. It is a stronger bond and cylindrically symmetrical.




## л-bond

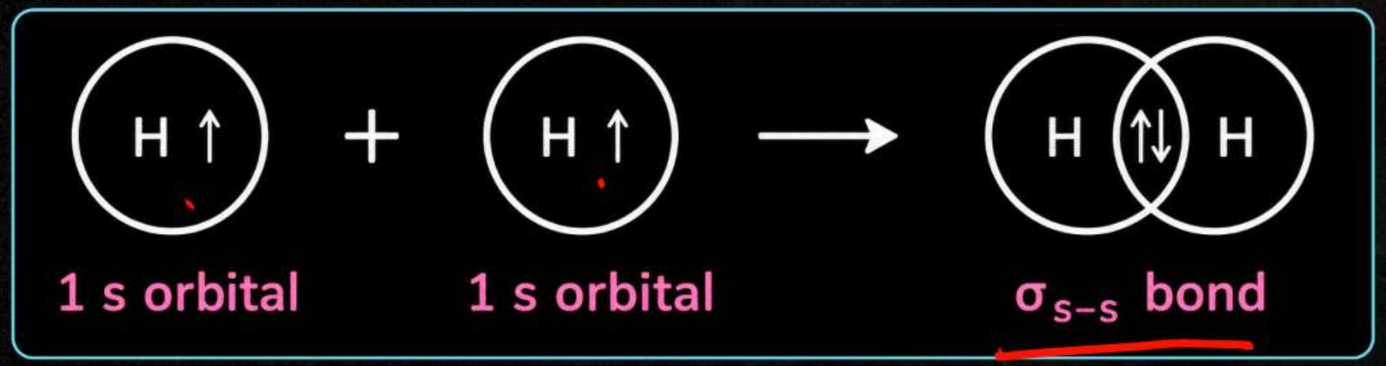
The covalent bond formed by sidewise overlapping of atomic orbital's is called л-bond. In this bond, the electron density is present above and below the inter nuclear axis.

It is relatively a weaker bond since the electrons are not strongly attracted by the nuclei ofbonding atoms.



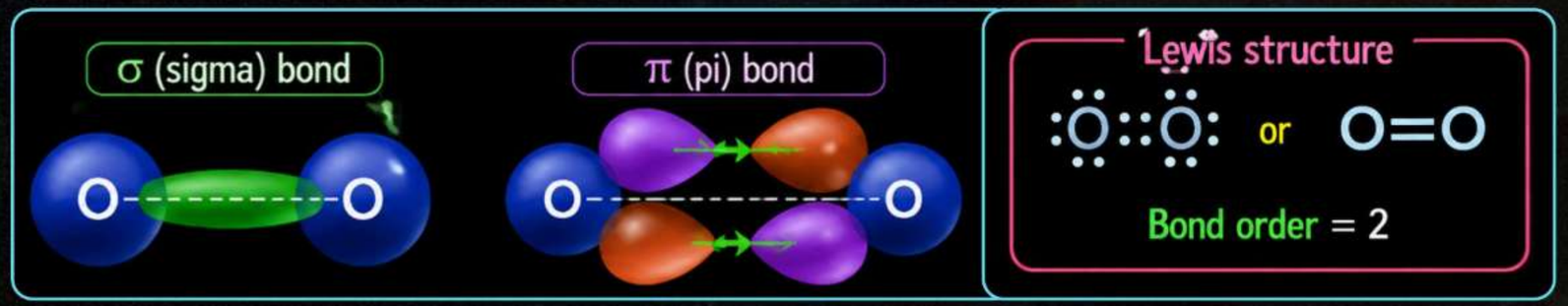
# $H_2$ molecule

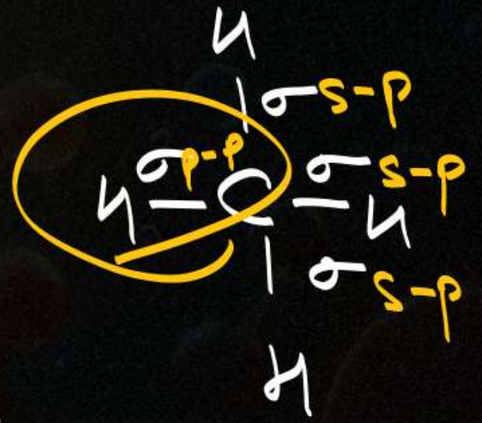
- The electronic configuration of hydrogen atom in the ground state is $1s^1$.
- In the formation of hydrogen molecule, two half filled 1s orbital's of hydrogen atoms overlap along the inter-nuclear axis and thus by forming a $\sigma_{s\text{-}s}$ bond.

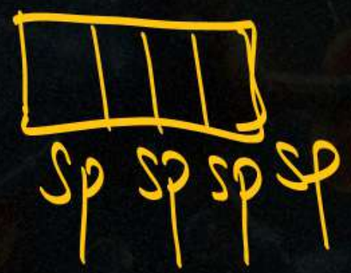


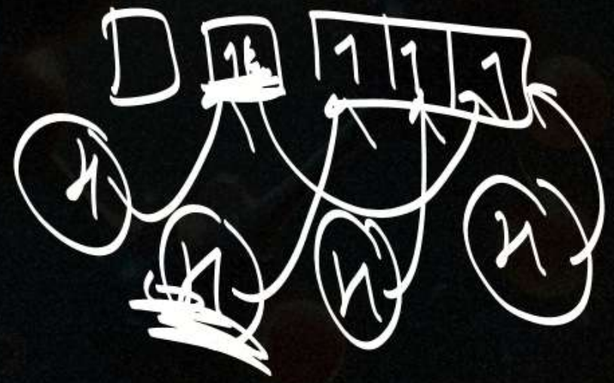
# $O_2$ molecule

- The electronic configuration of O in the ground state is [He] $2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2$.
- The half filled $2p_z$, orbital's of two oxygen atoms overlap along the inter-nuclear axis and form $\sigma_{p\text{-}p}$ bond.
- The remaining half filled $2p_z$ orbital's overlap laterally to form a $\pi_{p\text{-}p}$ bond.



# Hybridization

- The intermixing of two or more pure atomic orbital's of an atom with almost same energy to give same number of identical and degenerate new type of orbital's is known as hybridization.
- The new orbital's formed are also known as hybrid orbital's.
- During hybridization, the atomic orbital's with different characteristics are mixed with each other.
- Types of Hybridization - $sp$, $sp^2$, $sp^3$, $sp^3d$ , $sp^3d^2$

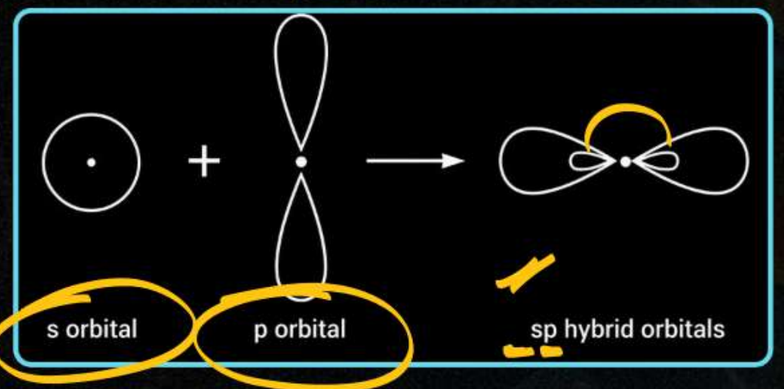


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Next Toppers

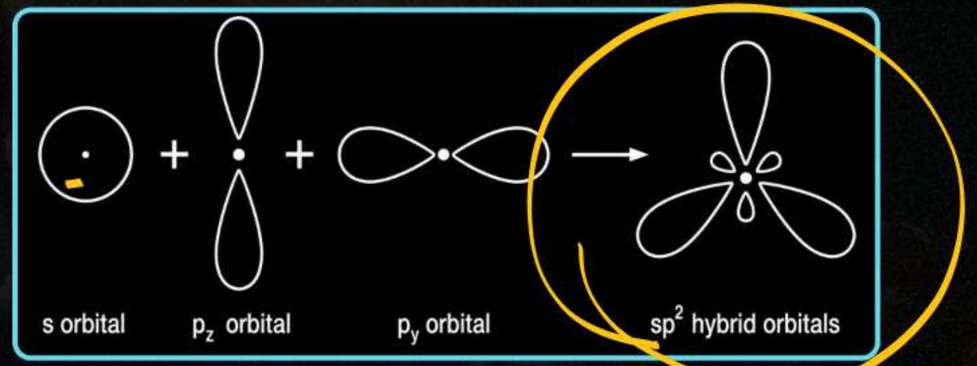
## sp Hybridization

- These sp-hybrid orbital's are arranged linearly at by making 180° of angle.
- They possess 50% 's' and 50% 'p'character.



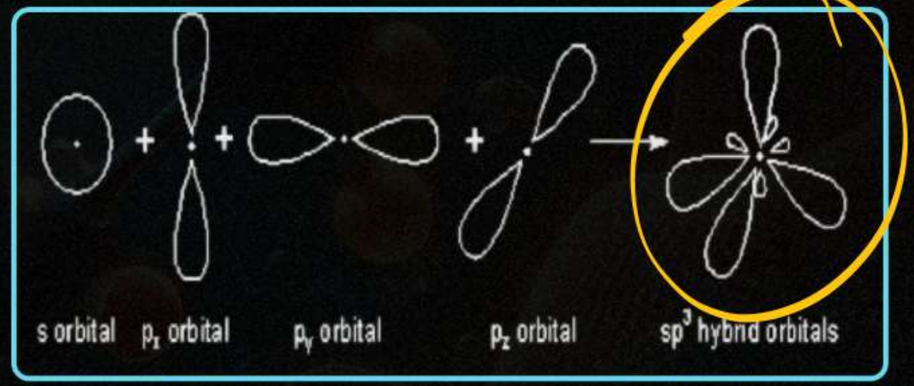

## $sp^2$ Hybridization

The three sp2 hybrid orbital's are oriented in trigonal planar symmetry at angles of 120° to each other.
They possess 33.3% 's'character and 66.6% 'p' character.



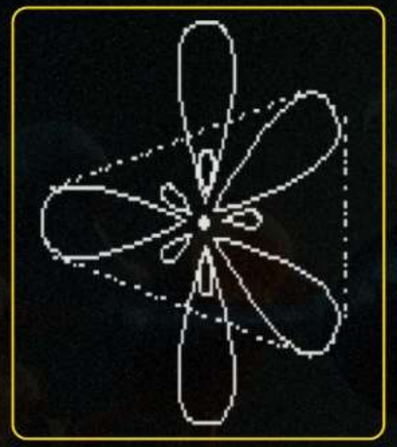

## $sp^3$ Hybridization

- These four sp3 hybrid orbital's are oriented in tetrahedral symmetry with 109° 28' angle with each other.
- They possess 25% 's' character and 75% 'p' character.



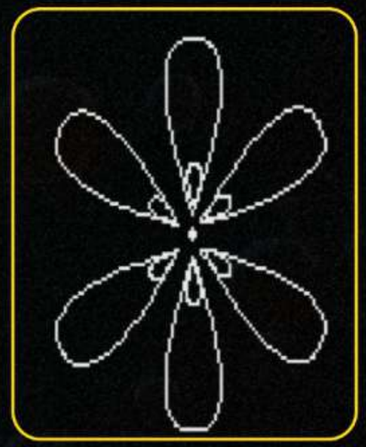
## $sp^3D$ Hybridization

- Arranged in trigonal bipyramidal symmetry.
- They possess 20% 's', 60% 'p' and 20% 'd' characters.



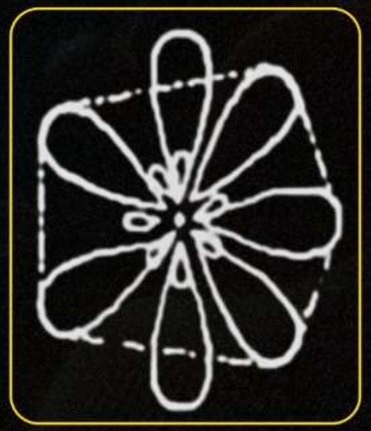
## $sp^3D^2$ Hybridization

- These six sp3d2 orbital's are arranged in octahedral symmetry by making 90° angles to each other.
- This arrangement can be visualized as four orbital's arranged in a square plane and the remaining two are oriented above and below this plane perpendicularly.

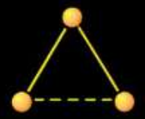
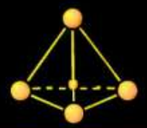
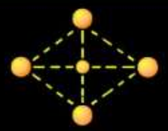
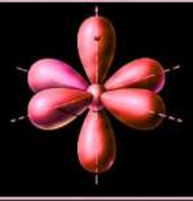


## $sp^3d^3$ Hybridization

- The sp3d3 hybrid orbital's, which are oriented in pentagonal bipyramidal symmetry.
- Five among the sp3d3 orbital's are arranged in a pentagonal plane by making 72° of angles. The remaining are arranged perpendicularly above and below this pentagonal plane.



# Hybridization

| Regions of Electron Density (Z) | Arrangement | Hybridization | |
|---|---|---|---|
| 2 |  linear | sp |  |
| 3 |  trigonal planar | $sp^2$ |  |
| 4 |  tetrahedral | $sp^3$ |  |
| 5 |  trigonal bipyramidal | $sp^3d$ |  |
| 6 |  octahedral | $sp^3d^2$ |  |

Z = 2 → sp

Z = 3 → $sp^2$

Z = 4 → $sp^3$

Z = 5 → $sp^3d$

Z = 6 → $sp^3d^2$

Z = $\sigma$ bond + L.p of C.A.

Z = ½ [valence $e^-$ + (-ve) - (+ve) + monovalent atom]

$H_2O \rightarrow z = \frac{1}{2}[6+2] \Rightarrow 4 \rightarrow sp^3 \rightarrow 104.5° \rightarrow$ bent




$NH_3 \rightarrow z = \frac{1}{2}[5+3] \Rightarrow 4 \rightarrow sp^3 \rightarrow 107° \rightarrow$ trigonal pyramidal

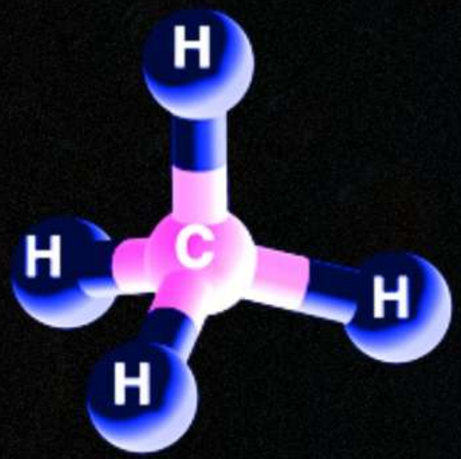
Diff shape



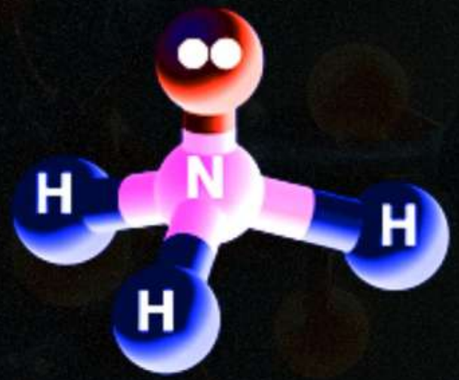

$CH_4 \rightarrow z = \frac{1}{2}[4+4] \Rightarrow 4 \rightarrow sp^3 \rightarrow 109°28 \rightarrow$ tetrahedral.



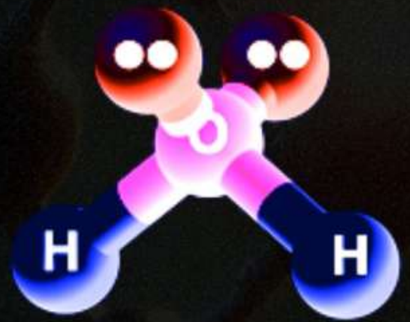
# Birth of VSEPR Theory




$AX_4E_0$
Bond Angle:
109.5°
Z= 4
$sp^3$




$AX_3E_1$
Bond Angle:
~107°
Z= 4
$sp^3$




$AX_2E_2$
Bond Angle:
~104.5°
Z= 4
$sp^3$

Same hybridisation but different bond angles ?

# VSEPR Theory

Shape → exclude L.p.

Geo → Include L.p.

L.p = 0

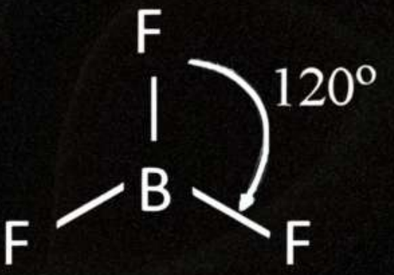
Shape = Geo

- The shape(excluding lone pairs) of a molecule depends upon the number of valence shell electron pairs (bonded or nonbonded) around the central atom.
- These pairs of electrons tend to occupy such positions in space that minimise repulsion and thus maximise distance between them. Lone pair - **Lone pair > Lone pair - Bond pair > Bond pair - Bond pair**
- A multiple bond is treated as if it is a single electron pair and the two or three electron pairs of a multiple bond are treated as a single super pair. **If lone pair is zero the shape = Geometry**

- There are two electron pairs in the valance shell of Beryllium. [$1s^2\ 2s^2$ ]
- Molecular geometry-Linear arrangement.




- Trigonal Planar Three electron pairs in the valance shell of Boron. [1s2 2s2 $2p^1$]
- Molecular geometry - Trigonal Planar arrangement.



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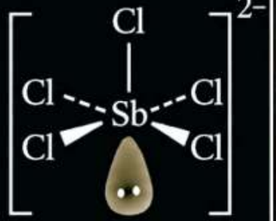
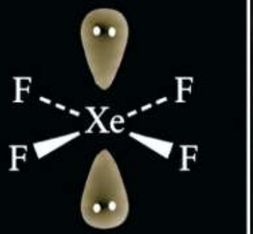
Next Toppers

| Number of Bonds | Number of Lone Pairs | Number of Charge Clouds | Molecular Geometry / Shape | Example |
|---|---|---|---|---|
| 2 | 0 | 2 | Linear | O=C=O |
| 3 | 0 | 3 | Trigonal planar | H₂C=O |
| 2 | 1 | | Bent | SO₂ |
| 4 | 0 | 4 | Tetrahedral | CH₄ |
| 3 | 1 | | Trigonal pyramidal | NH₃ |
| 2 | 2 | | Bent | H₂O |

| Number of Bonds | Number of Lone Pairs | Number of Charge Clouds | Molecular Geometry / Shape | Example |
|---|---|---|---|---|
| 5 | 0 | 5 | Trigonal bipyramidal | PCl₅ |
| 4 | 1 | | Seesaw | SF₄ |
| 3 | 2 | | T-shaped | ClF₃ |
| 2 | 3 | | Linear | $I_3^-$ |

BACKLOG MARATHON SERIES

Next Toppers

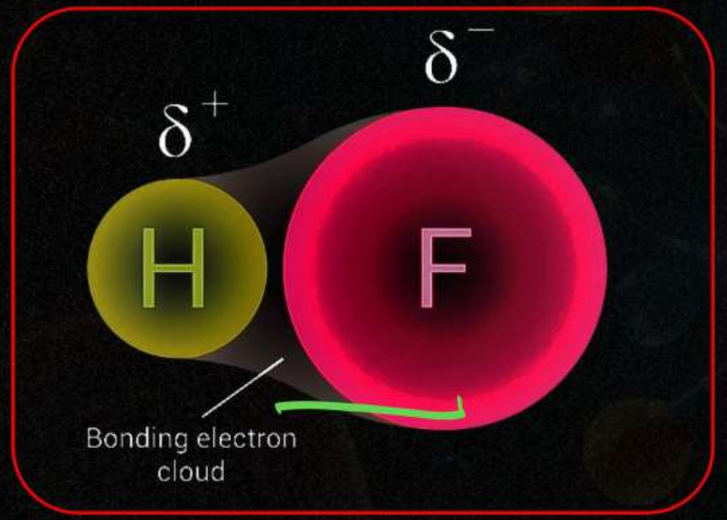
| Number of Bonds | Number of Lone Pairs | Number of Charge Clouds | Molecular Geometry | Example |
|---|---|---|---|---|
| 6 | 0 | 6 | Octahedral |  |
| 5 | 1 | | Square pyramidal |  |
| 4 | 2 | | Square planar |  |

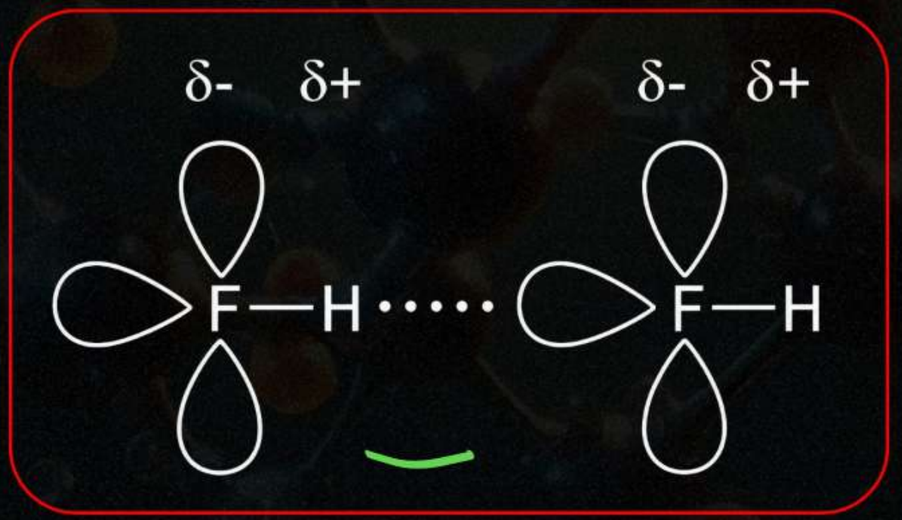
# Limitations of VSEPR

- It fails to predict the shapes of isoelectronic species and transition metal compounds.
- Not based on quantum theory.
- This model does not take relative sizes of substituents.
- Unable to explain atomic orbitals overlap.

# Hydrogen Bonding

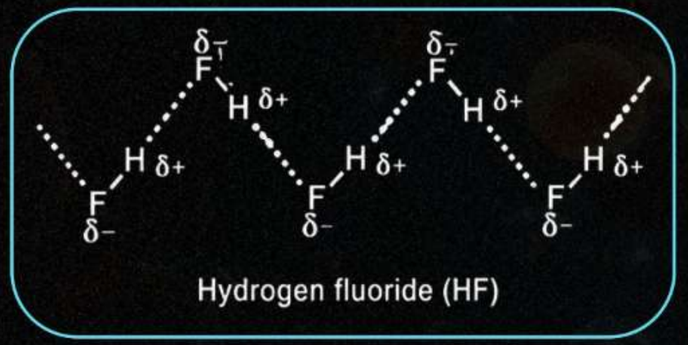
- When hydrogen is bonded to a highly electronegative atom (such as nitrogen, oxygen, fluorine), the bonding electron pair is drawn towards the electronegative atom.
- Hydrogen has no inner shell electron and is very small in size, the positive charge density eveloped is high.
- Electrostatic attraction between hydrogen atoms bonded to small, strongly electronegative atoms (N, O and F) and the lone pair electrons on these electronegative atoms.

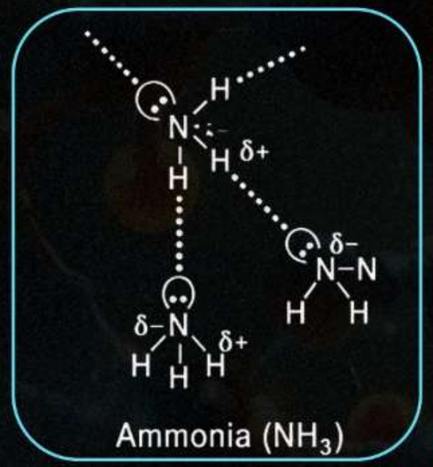


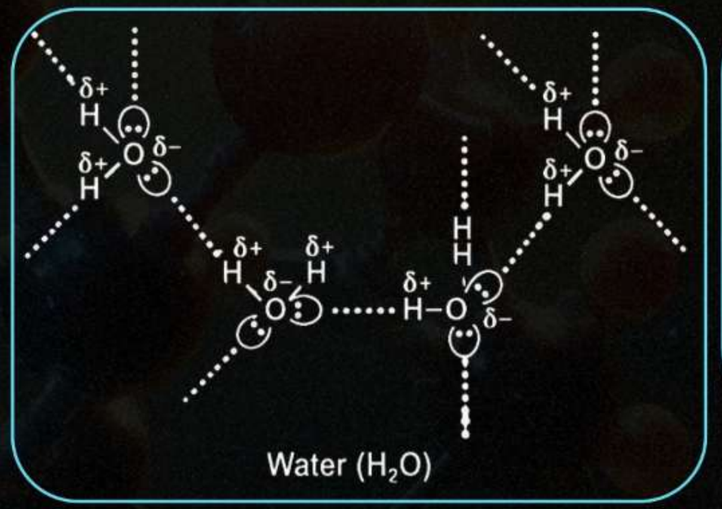



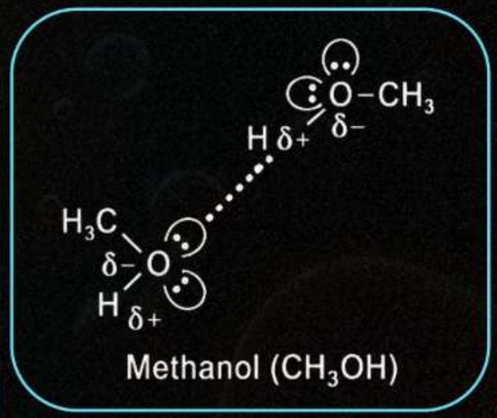
# Essential Requirements for Formation of Hydrogen Bond

- 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.



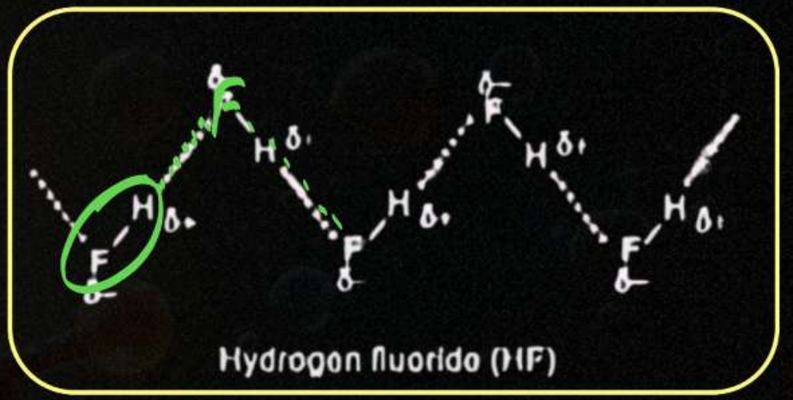



# Types of Hydrogen Bond

## 1. Intermolecular hydrogen bond

- Intermolecular hydrogen bonding is a type of weak chemical bond that occurs between molecules.
- It happens when a hydrogen atom, already covalently bonded to a highly electronegative atom (like oxygen, nitrogen, or fluorine), is attracted to another nearby electronegative atom in a different molecule.



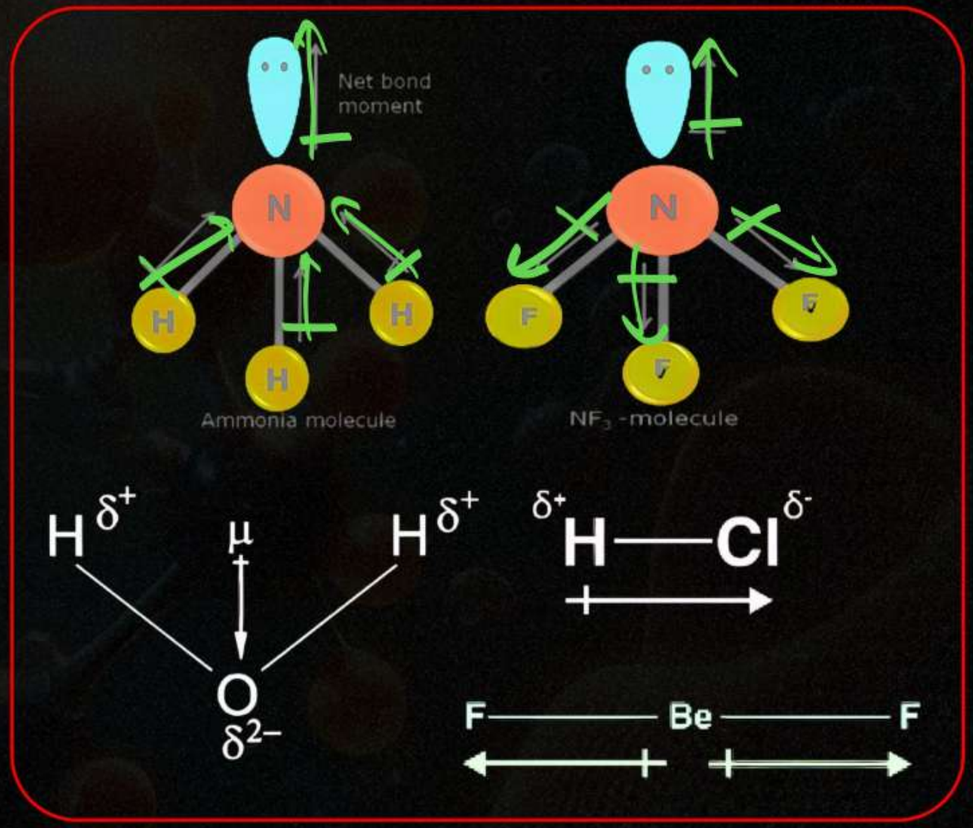

## 2. Intramolecular hydrogen bond

- Intramolecular hydrogen bonding is a type of hydrogen bond that occurs within a single molecule, specifically when a hydrogen atom is simultaneously bonded to two highly electronegative atoms (like oxygen, nitrogen, or fluorine) within the same molecule.
- This creates a ring-like structure or a chelation within the molecule.



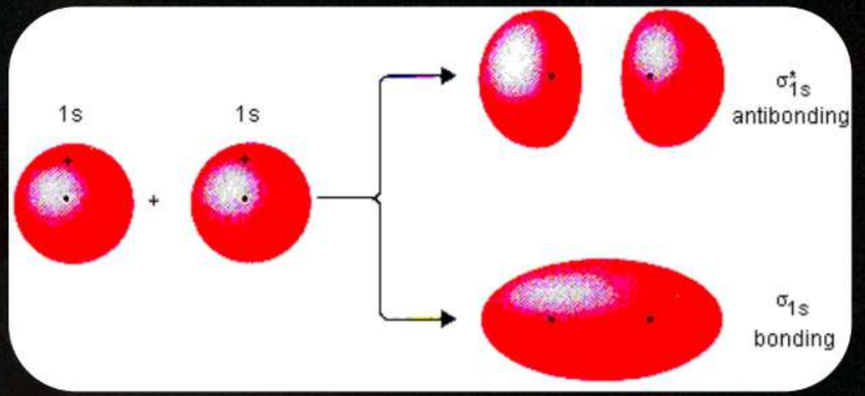
# Dipole Moment

- Dipole moment (μ) is the product of the magnitude of the charge (Q) and the distance (r) between the charges.
- It's a vector quantity, meaning it has both magnitude (strength) and direction. Its direction is from less E.N. to more E.N. element or from atom to lone pair.
- If net dipole moment = 0, it is a non polar molecule.
- Dipole Moment of homoatomic molecule is zero e.g. $H_2$,$Cl_2$,$Br_2$ etc.
- $NH_3$ has a significantly higher dipole moment (approx. 1.47 D) than $NF_3$ (approx. 0.24 D). This is because in $NH_3$, the bond dipoles of the N-H bonds and the lone pair on nitrogen all align in the same general direction, reinforcing each other, while in $NF_3$, the highly electronegative fluorine atoms pull electrons from the nitrogen atom, and the N-F bond dipoles oppose the dipole of the lone pair, causing a partial cancellation.



# Molecular Orbital Theory

- Allows one to predict the distribution of electrons and this in turn can help predict molecular properties such as shape, magnetism, and Bond Order.




## Salient Features of M.O.T

(i) The electrons in a molecule are present in the various molecular orbitals as the electrons of atoms are present in the various atomic orbitals.

(ii) The atomic orbitals of comparable energies and proper symmetry combine to form molecular orbitals.

(iii) While an electron in an atomic orbital is influenced by one nucleus, in a molecular orbital it is influenced by two or more nuclei depending upon the number of atoms in the molecule. Thus,an atomic orbital is monocentric while a molecular orbital is polycentric.

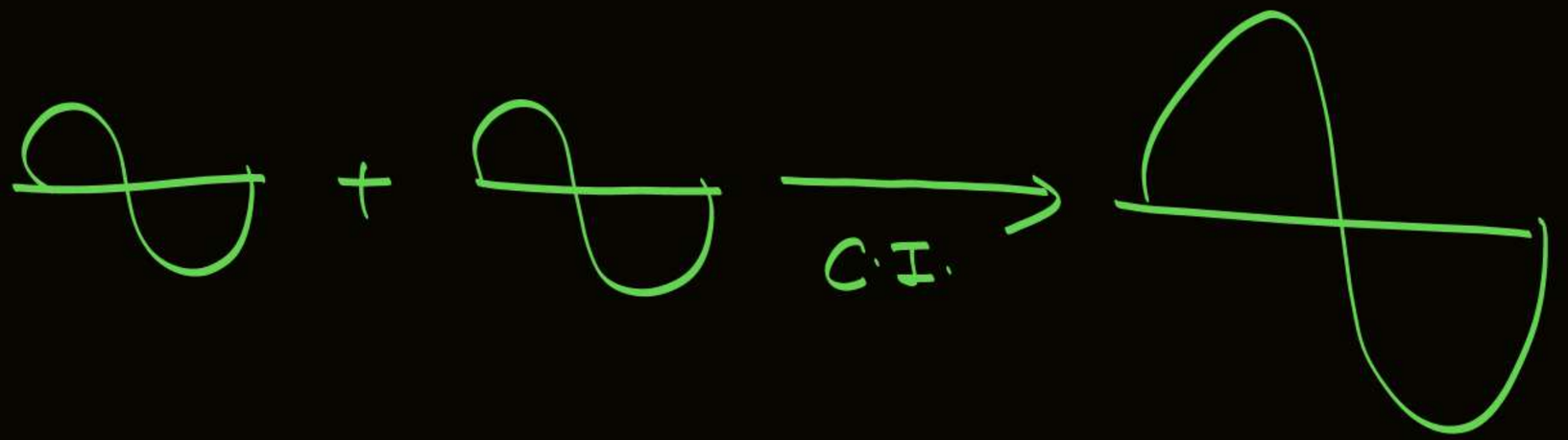
# Salient Features of M.O.T

(iv) The number of molecular orbital formed is equal to the number of combining atomic orbitals. When two atomic orbitals combine, two molecular orbitals are formed. One is known as bonding molecular orbital while the other is called antibonding molecular orbital.

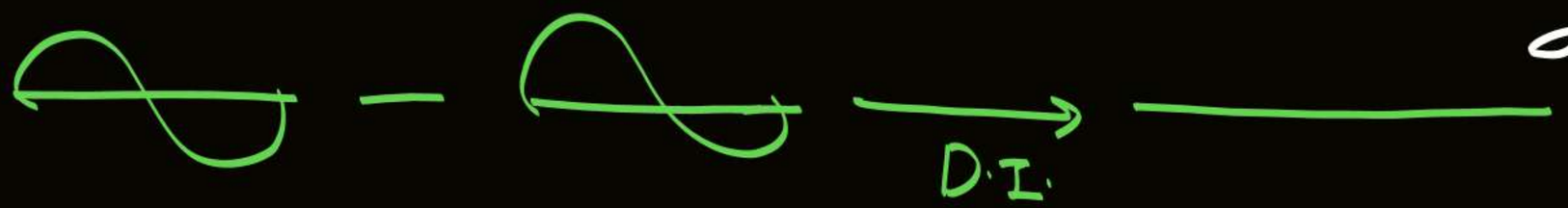
(v) The bonding molecular orbital has lower energy and hence greater stability than the corresponding antibonding molecular orbital.

(vi) Just as the electron probability distribution around a nucleus in an atom is given by an atomic orbital, the electron probability distribution around a group of nuclei in a molecule is given by a molecular orbital.

(vii) The molecular orbitals like atomic orbitals are filled in accordance with the aufbau principle obeying the Pauli's exclusion principle and the Hund's rule.



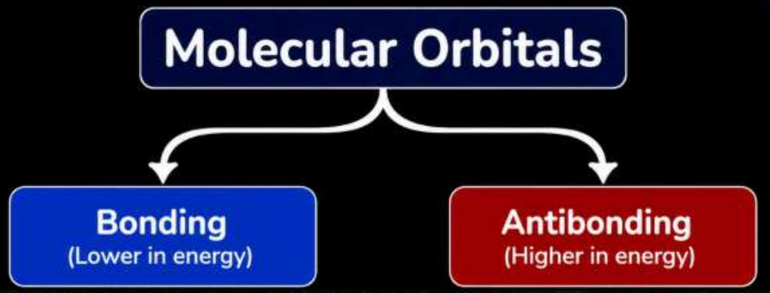

$\sigma$ Energy ↓ Stability ↑
Bonding orbital

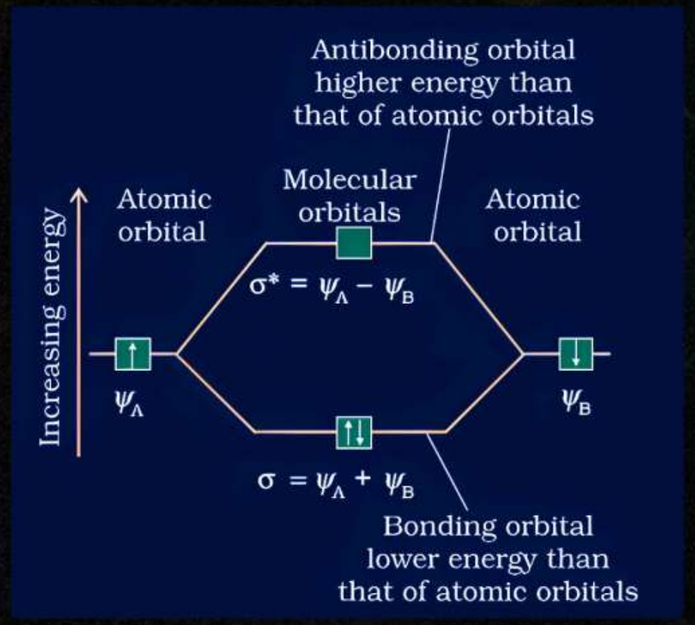


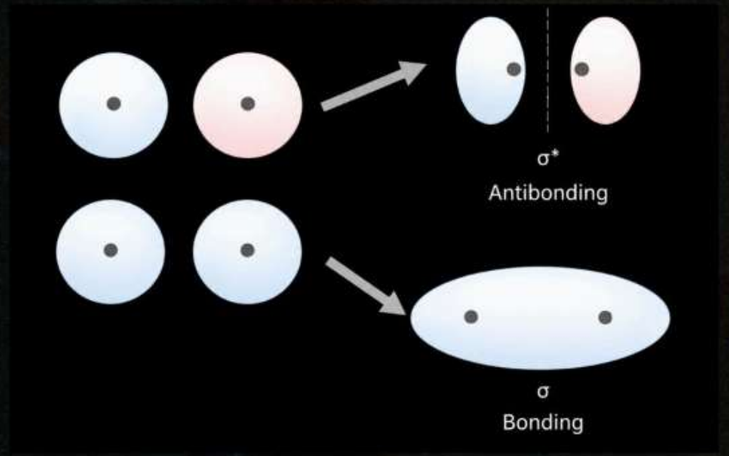

$\sigma^*$ Energy ↑ stability ↓
Antibonding

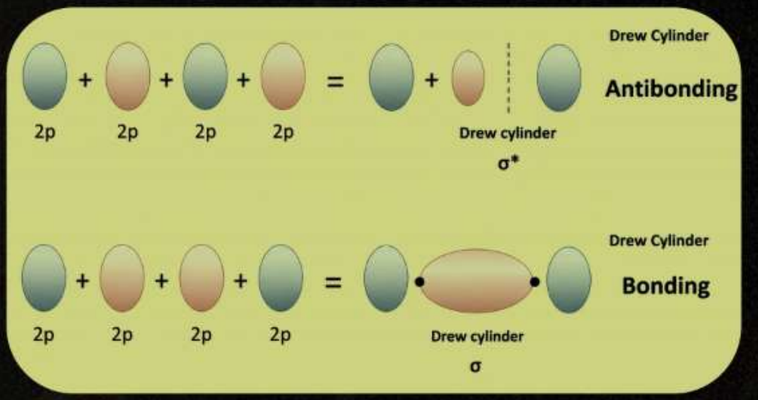
# Principles of MOT

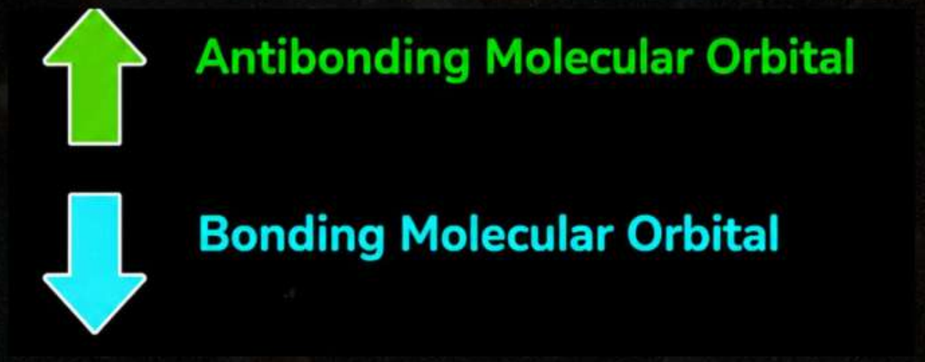
**In molecules, atomic orbitals combine to form molecular orbitals which surround the molecule.**



# E.C. Molecular Orbital

($e \leq 14$)

$$\sigma 1s \;\; \sigma^* 1s \;\; \sigma 2s \;\; \sigma^* 2s \;\; (\pi 2p_x = \pi 2p_y) \;\; \sigma 2p_z \;\; (\pi^* 2p_x = \pi^* 2p_y) \;\; \sigma^* 2p_z$$

($e > 14$)

$$\sigma 1s \;\; \sigma^* 1s \;\; \sigma 2s \;\; \sigma^* 2s \;\; \sigma 2p_z \;\; (\pi 2p_x = \pi 2p_y) \;\; (\pi^* 2p_x = \pi^* 2p_y) \;\; \sigma^* 2p_z$$

$C_2$ → 12 $e \leq M$

E.C. → $\sigma 1s^2 \; \sigma^* 1s^2 \; \sigma 2s^2 \; \sigma^* 2s^2 \; (\pi 2p_x^2 = \pi 2p_y^2)$

B.O → $\frac{N_b - N_a}{2} = \frac{8-4}{2} = \frac{4}{2} = 2$ → C = C

Magnetic B → Diamagnetic.

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# Electronic Configuration in MOT

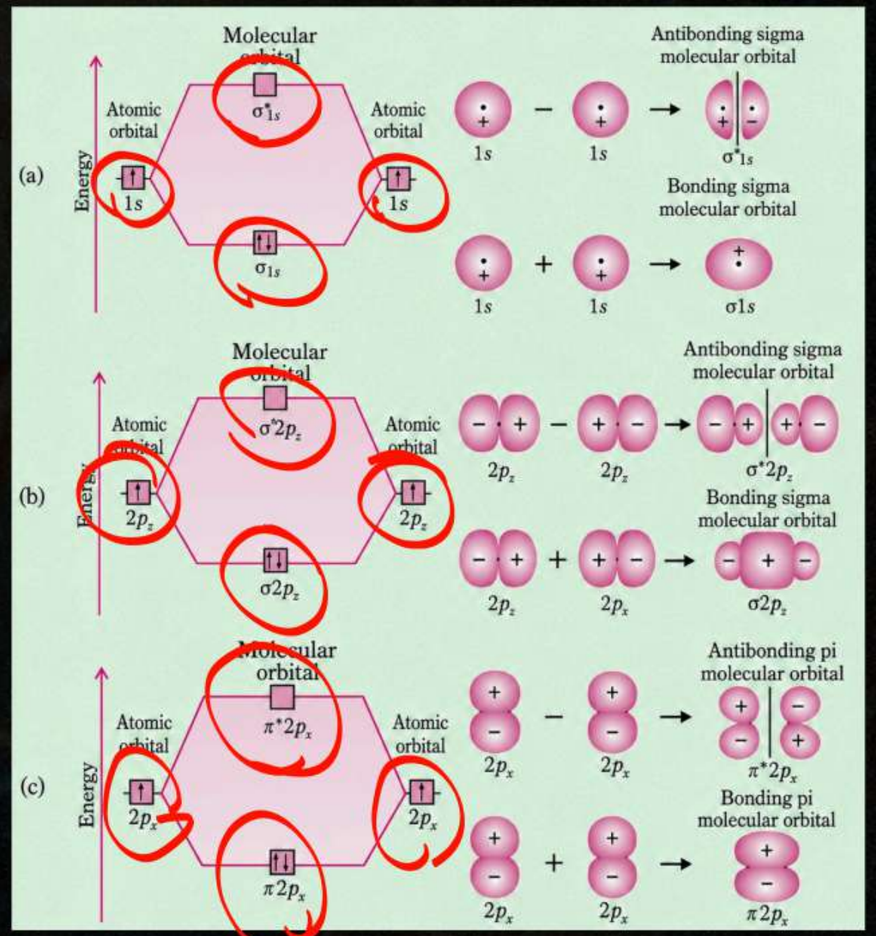
## Electrons ≤ 14

$$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x = \pi 2p_y) < \sigma 2p_z < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$$

## Electrons > 14

$$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \sigma 2p_z < (\pi 2p_x = \pi 2p_y) < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$$

## Energy Level diagram for molecular orbitals



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# Electronic Configuration in MOT

$O_2$: $(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma^* 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2$

$\left(\pi 2p_x^{\ 2} \equiv \pi 2p_y^{\ 2}\right) \left(\pi^* 2p_x^{\ 1} \equiv \pi^* 2p_y^{\ 1}\right)$

$$O_2 : \begin{bmatrix} KK\ (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 \\ \left(\pi 2p_x^2 \equiv \pi 2p_y^2\right), \left(\pi^* 2p_x^1 \equiv \pi^* 2p_y^1\right) \end{bmatrix}$$

$C_2 : (\sigma 1s)^2\ (\sigma^* 1s)^2\ (\sigma^* 2s)^2\ (\pi 2p_x^2 = \pi 2p_y^2)$

or $KK\ (\sigma 2s)^2\ (\sigma^* 2s)^2\ (\pi 2p_x^2 = \pi 2p_y^2)$

The bond order of $C_2$ is $\frac{1}{2}\ (8 - 4) = 2$

# Electronic Configuration and Molecular Behaviour

## 1. Stability of Molecules

If Nb is the number of electrons occupying bonding orbitals and Na the number occupying the antibonding orbitals, then,

The molecule is stable if Nb is greater than Na, and

The molecule is unstable if Na is less than Nb

## 2. Bond order

$\frac{N_b - N_a}{2}$

Bond order (b.o.) is defined as one half the difference between the number of electrons present in the bonding and the antibonding orbitals

## 3. Bond Length

The bond order between two atoms in a molecule may be taken as an approximate measure of the bond length. The bond length decreases as bond order increases.

## 4. Magnetic Nature

If all the molecular orbitals in a molecule are doubly occupied, the substance is diamagnetic (repelled by magnetic field). However if one or more molecular orbitals are singly occupied it is paramagnetic (attracted by magnetic field), e.g. O2 molecule.


Khatam !
Tata !!
Bye bye !!!
Fir Mileinge
prarambh mein
Killiar

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