EKADEMY

Structure of Atom

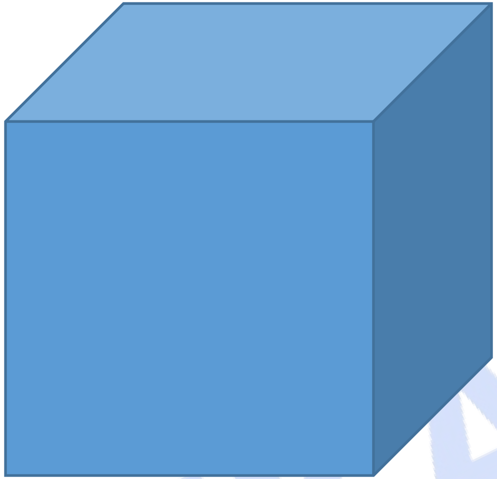
Chapter 2

Class 11

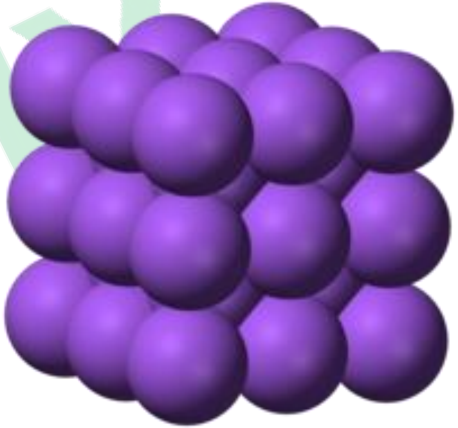
Abhinav K Singh (AKS)

Chapter 2
Structure of Atom

Introduction of Atom



made of →



Matter

Atoms → Greek word → *atomio* → Indivisible

Atoms → Smallest Particle → Indivisible

EKADEMY

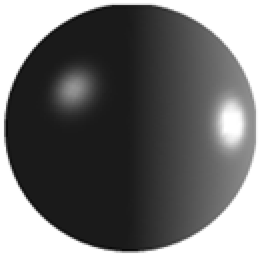
Dalton's Atomic Theory

John Dalton

Law of conservation of mass

Law of constant composition

Law of multiple proportion

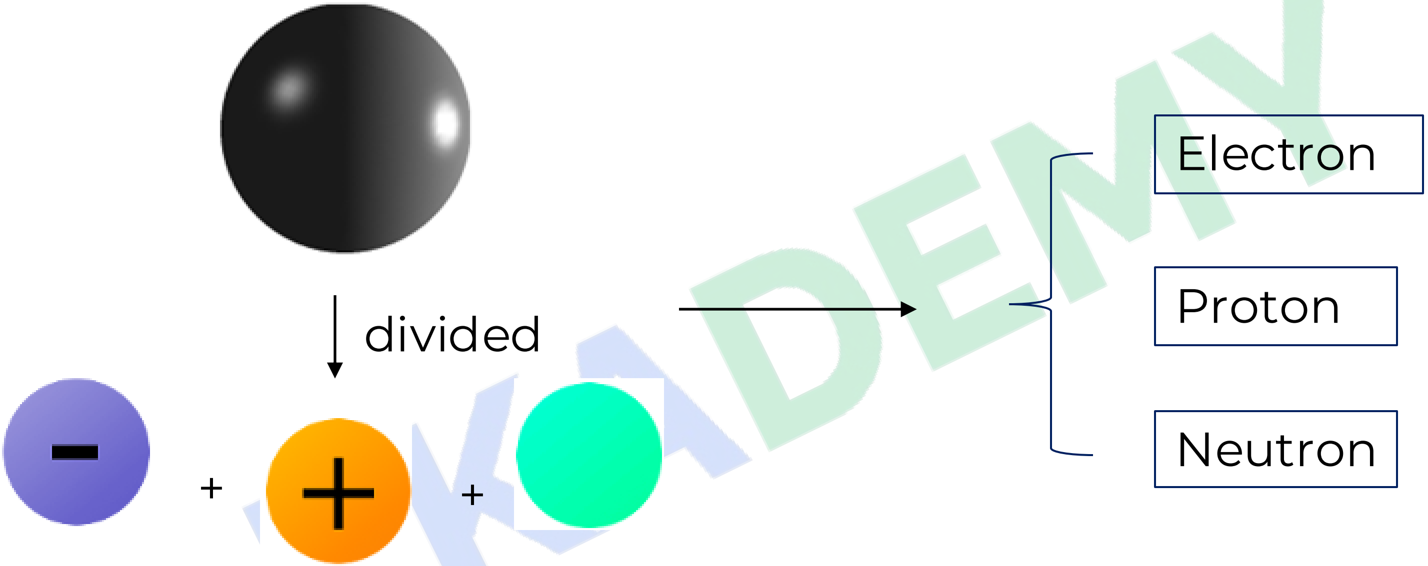


Neutral Atom

Unable to explain:

Generation of electrical charge while rubbing glass by silk

Daltons Assumption that atoms are indivisible was wrong



2.1

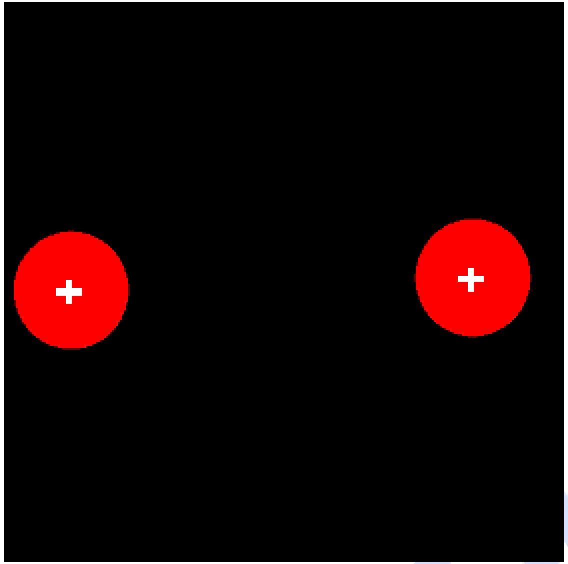
Discovery of Sub-atomic Particles

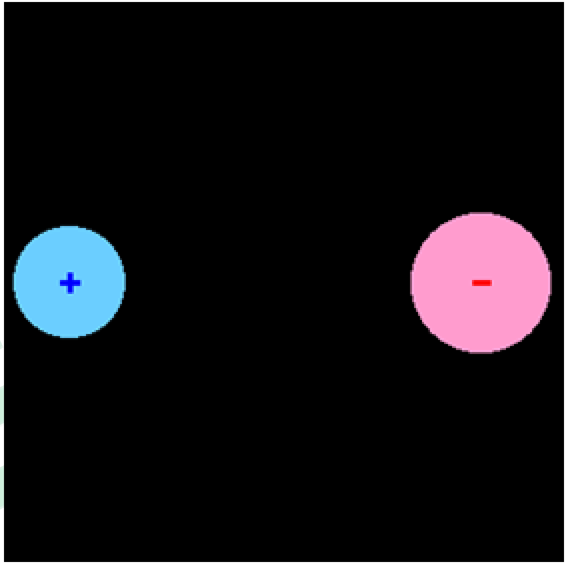
Discovery of Subatomic Particles

Learning Objectives

- Discovery of Electron
- Discovery of Proton & Neutron

1. Discovery of Electron

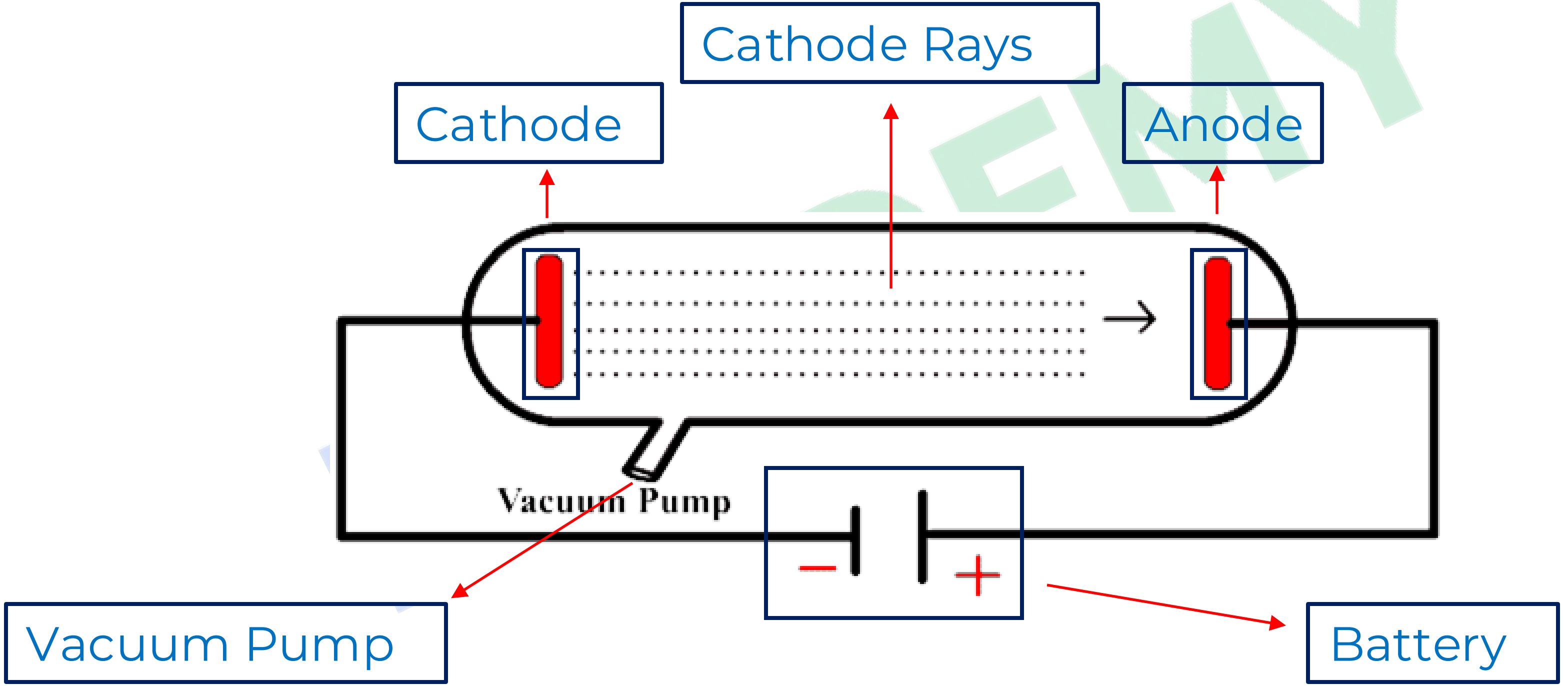




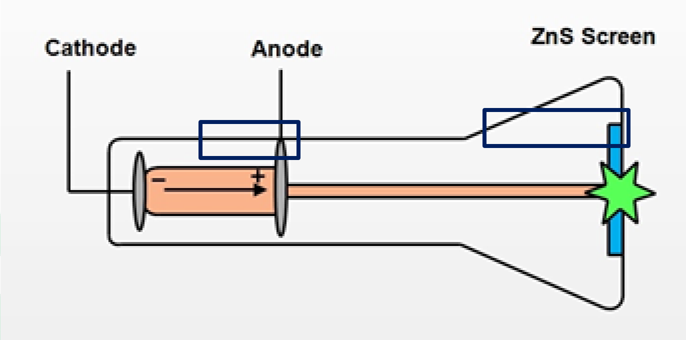
Like charges repel each other

Unlike charges attract each other

Experimental Setup: Cathode Ray Discharge Tube



J.J. Thomson



Modifications:

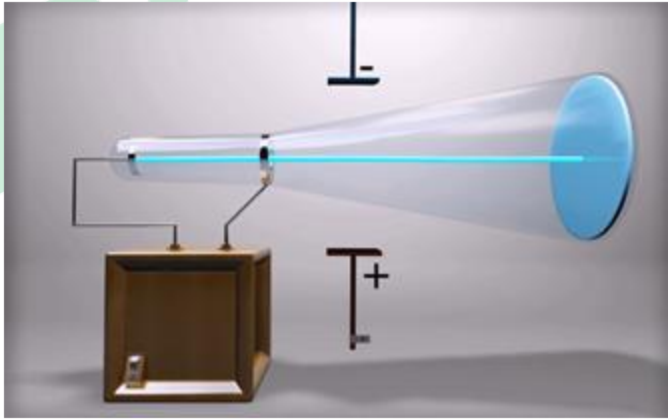
1. Completely evacuated tube

2. Hole was created in Anode

3. ZnS screen was placed behind the Anode

Result of these Experiment:

1. Cathode rays move from cathode to Anode
2. Cathode rays are invisible
3. These rays travel in straight line
4. In presence of electric field and magnetic field, it deflected towards positive plate
5. Characteristics of CR is independent of Material of Electrodes and Nature of Gas

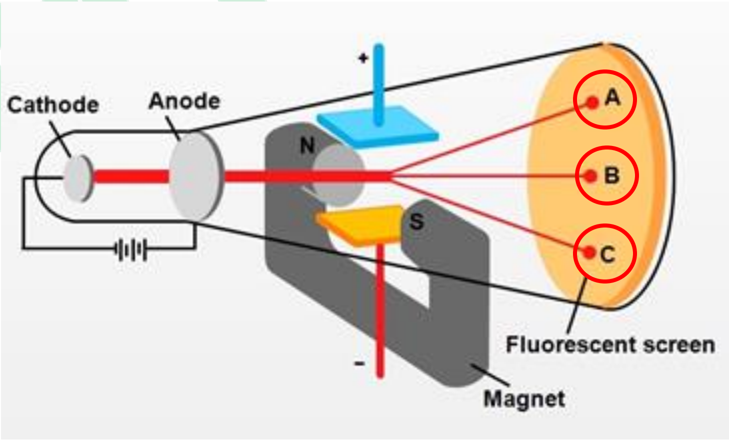


Cathode Rays ⟶ Negative charge particle ⟶ Electron

J.J. Thomson

Deflection depends on:

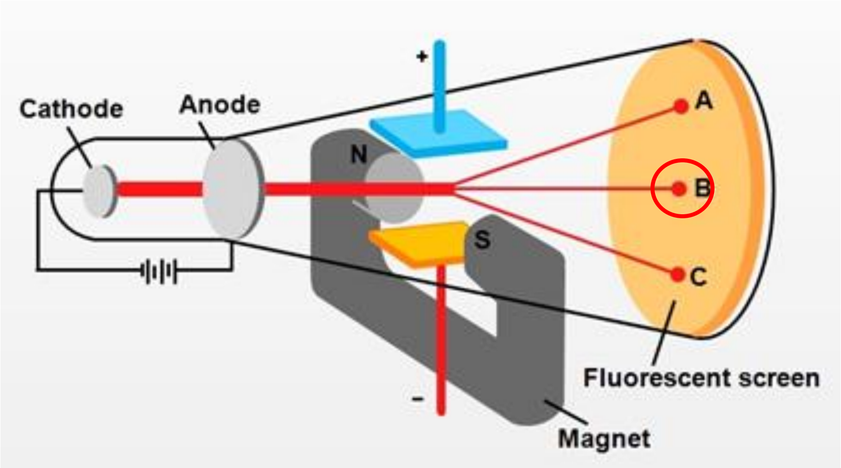
1. Mass of electron
2. Charge of electron
3. Strength of the Field



Electric Force

Static

Magnetic Force

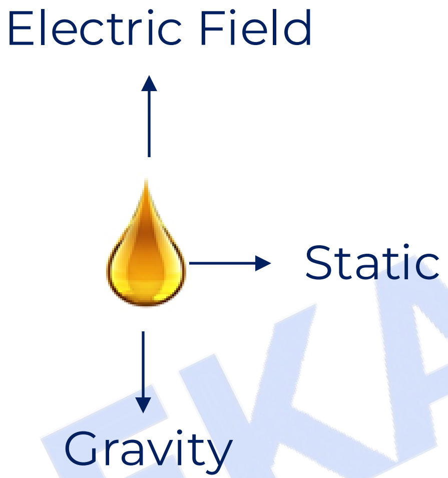


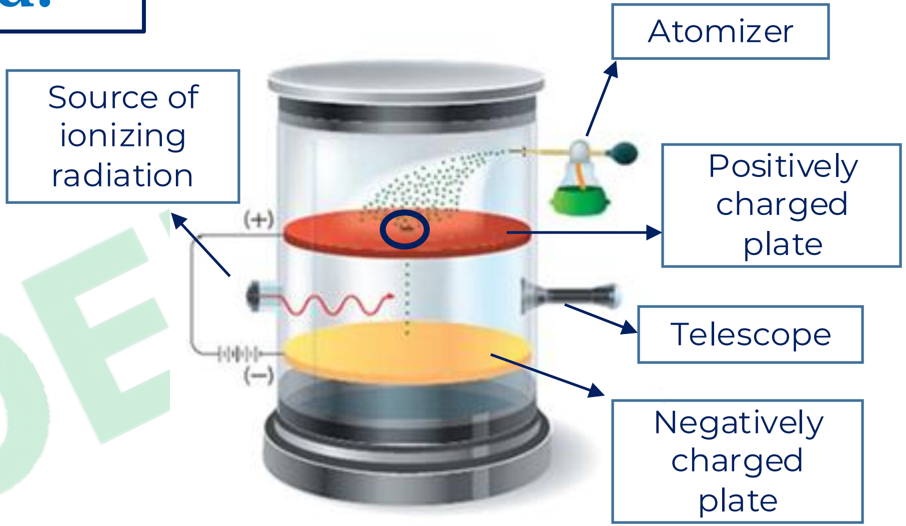
$$\boxed{\frac{e}{m_e} = 1.758820 \times 10^{11}\ \mathrm{C\ Kg^{-1}}}$$

e = charge of electron(−ve)

$\mathrm{m_e}$ = mass of electron

Millikan's oil drop Method:





Millikan's oil drop Apparatus

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

Millikan's oil drop Method:

$$\boxed{e = -1.6 \times 10^{-19}\, C}$$

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

$$m_e = \frac{e}{e/m_e} = \frac{-1.602176 \times 10^{-19}\, C}{1.758820 \times 10^{11}\, CKg^{-1}}$$

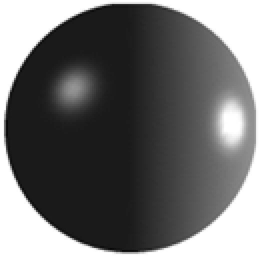
$$\boxed{m_e = 9.1094 \times 10^{-31}\, Kg}$$

2. Discovery of Proton & Neutron

Discovery of Proton

Modification:

1. CRT is not evacuated & filled with gas
2. Hole was created in Cathode
3. ZnS screen was placed behind the Cathode



Atom

↓

Neutral

↓

Positive + Negative

Discovery of Proton

Observations:

1. Glow was seen: indicate charged particle

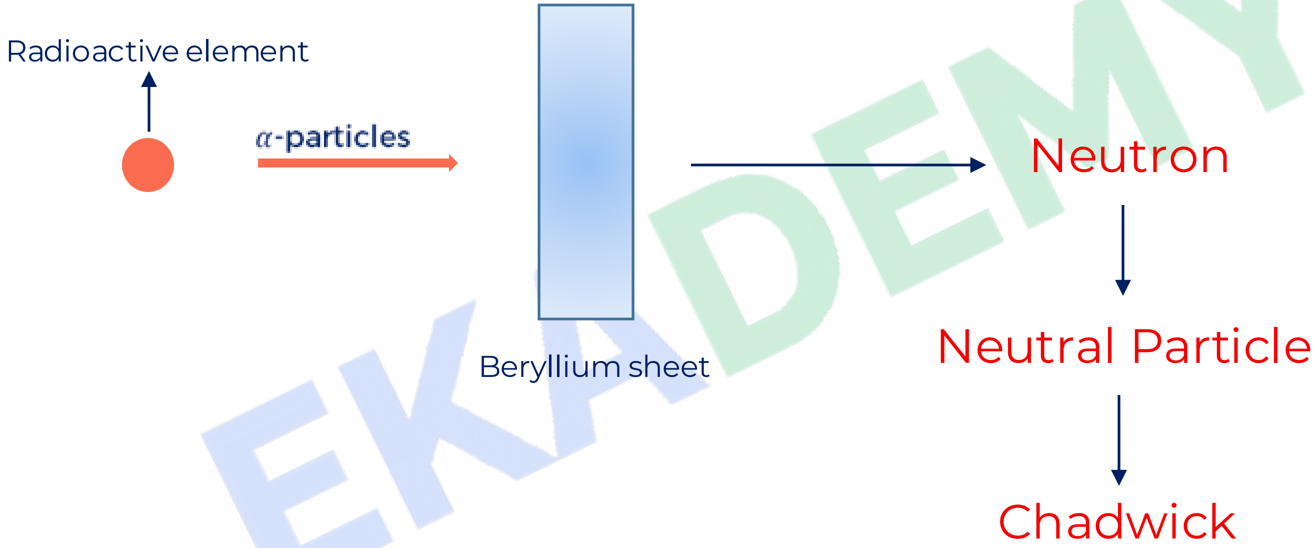
2. $\frac{q}{m}$ depends upon Gas

3. In presence of electric field, it deflects toward negative plate: Indicate positively charged particle

4. Mass of Particle depends on Nature of Gas

Smallest & Lightest positive ion $\longrightarrow$ Hydrogen $\longrightarrow$ Proton

Discovery of Neutron



Properties of Sub-atomic Particles:

Name	Symbol	Absolute charge/C	Relative charge	Mass/kg	Mass/u	Approx. mass/u
Electron						
Proton						
Neutron						

2.1

Solved problems

Q. What is the decreasing order of e/m ratio of electron, proton and neutron?

Sol:

Electron		Proton		Neutron
$\frac{e}{m_e}$	>	$\frac{e}{\approx 1836\ m_e}$	>	$\frac{0}{m_e} = 0$

Electron > Proton > Neutron

Summary

- Atoms are divisible .
- Sub atomic particles are electron, proton and neutron .
- Discovery of Electron : Mass $= 9.109382 \times 10^{-31}$ kg ,
  Charge $= -1.602176 \times 10^{-19}$ C
- Discovery of Proton : Mass $= 1.6726216 \times 10^{-27}$ Kg
  Charge $= +1.602176 \times 10^{-19}$ C
- Discovery of Neutron : Mass $= 1.674927 \times 10^{-27}$
  Charge $= 0$ C

2.1 Discovery of Subatomic Particles

Reference Questions

NCERT Exercises: 1

2.2

Atomic Models

2.2 Atomic Models

Learning Objectives

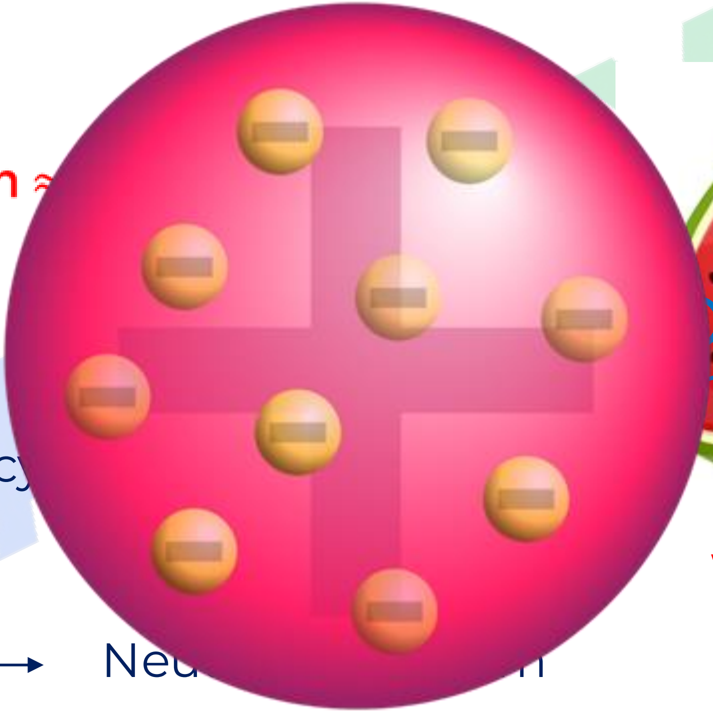
- Thomson Model of Atom & Rutherford Model
- Atomic number , Mass number , Isotopes, Isobars & Isotones

2.2.1 Thomson Model of Atom & Rutherford Model

Thomson Model of Atom

J.J. Thomson

Atom



Red juic

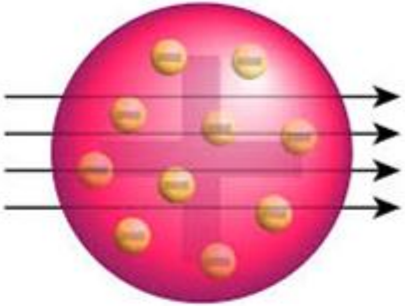
Watermelon

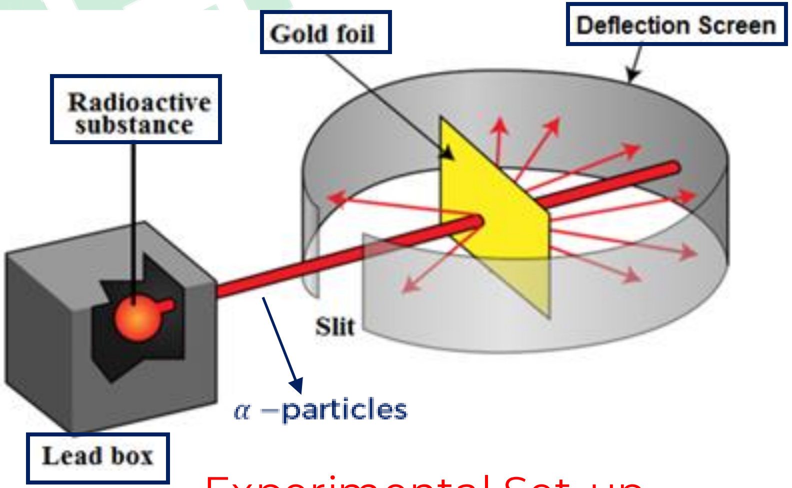
Explained: → Neu

Atom consists of Positively charged sphere with uniform distribution of electrons

Rutherford's Model

Rutherford



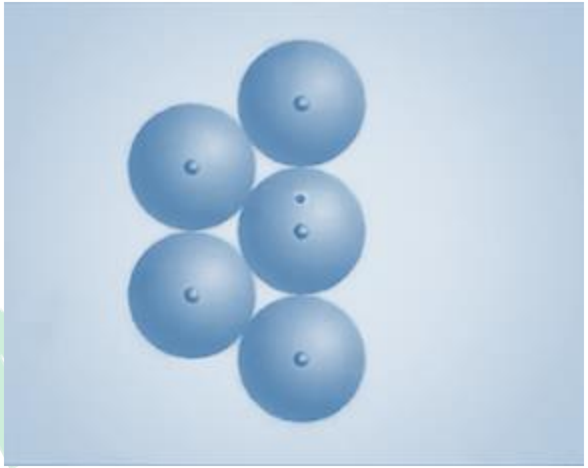


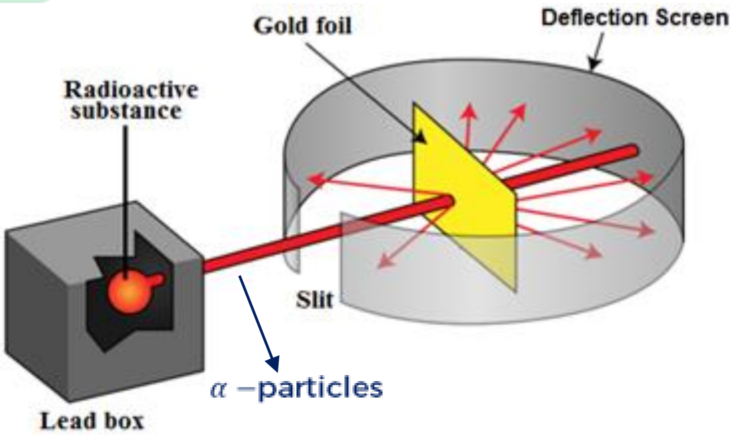
Experimental Set-up

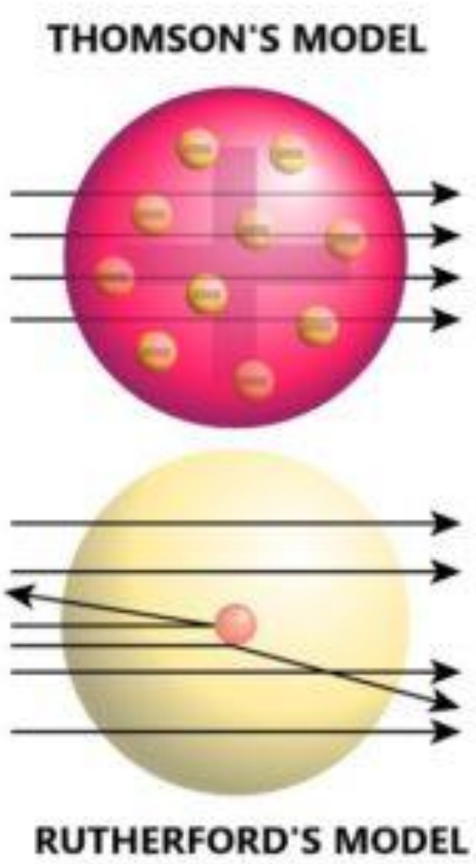
Rutherford's Model

Observations:

- Most of the α – particles went undeflected.
- Few α – particles deflected by small angle.
- Very few α – particles deflected backwards.

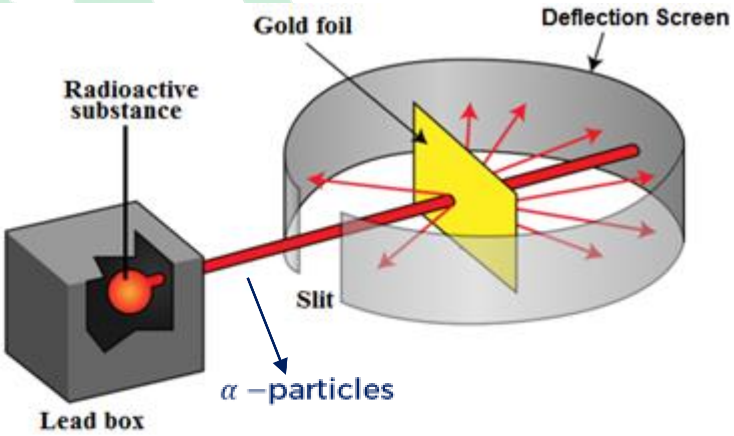





THOMSON'S MODEL
RUTHERFORD'S MODEL

Conclusions drawn by Rutherford:

1. Atoms have lot of empty space
2. Atom have positive charged entity
3. Positive charged entity is present as dense mass at the center
4. Dense mass is very small & is known as Nucleus



Size of Nucleus = 10^{-15} m

Size of Atom = 10^{-10} m

Rutherford's Model

Nuclear Model of Atom:

1. Positively charged mass → Located in centre → Nucleus
2. Electrons → Circular Path → Orbit
3. Force → Electrostatic force → Held Together

Limitations:

Rutherford's model could not explain

- Stability of electron in an orbit
- Distribution of electrons
- Atomic Spectra

2.2.2 Atomic Number & Atomic Mass

Atomic Number

z = No. of Protons

Atomic Mass

A = No. of Protons + No. of Neutrons

Representation of Element

Element Symbol

$^{A}_{Z}X$ $^{4}_{2}He$

No. of Electrons in Neutral Atom

No. of electron = Z

$^{4}_{2}He$

Z = 2 = No. of electrons

Neutral Atom

Positive Charge → Protons

Negative Charge → Electrons

Protons = Electrons

Charged Species

Positively charged	Negatively charged

$${}^{24}_{12}Mg^{+2}$$

No. of Protons > No. of electrons

e = 12-2 = 10

$${}^{14}_{7}N^{-3}$$

No. of Protons < No. of electrons

e = 7+3 = 10

Isotopes, Isotones & Isobars

Isotopes

Equal no of protons

$^{35}_{17}Cl$ $^{37}_{17}Cl$

$\boxed{17}$

Isotones

Equal no of Neutrons

$^{23}_{11}Na$ $^{24}_{12}Mg$

$\boxed{12}$

Isobars

Same atomic mass

$^{14}_{6}C$ $^{14}_{7}N$

$\boxed{14}$

Solved Problems

Q. Two nuclides A and B are isoneutronic. Their atomic mass are 76 and 77 respectively. If the atomic number of A is 32 , find the atomic number of B .

Sol: Given: A & B are isoneutronic, $A_A = 76$ $A_B = 77$, $Z_A = 32$, $Z_B = ?$

Neutrons in A = Neutrons in B

Neutrons = Atomic mass – Atomic Number

$$A_A - Z_A = A_B - Z_B$$

$$76 - 32 = 77 - Z_B$$

$$Z_B = 33$$

Concept Test

Q. Calculate the number of electrons which will together weigh one gram.

Q. Calculate the number of electrons which will together weigh one gram.

Sol:

Mass of 1 electron $= 9.1 \times 10^{-31}$ kg

$= 9.1 \times 10^{-28}$ g

Number of electrons that weighs $1\ g = \frac{1}{9.1 \times 10^{-28}}$

$\boxed{= 1.098 \times 10^{27} \text{ electrons}}$

Summary

- Thomson Model = Watermelon Model
- **Rutherford Model : α- particle scattering experiment**
- Atomic Number = Number of protons in the nucleus of an Atom
  = Number of electrons in a neutral Atom
- Mass Number = Number of protons + Number of neutrons
- Isotopes = Equal number of protons
- Isotones = Equal number of neutrons
- Isobars = Same atomic mass

Atomic Models

Reference Questions

NCERT Exercises: 2, 3, 4, 22, 26(i), 27, 42, 43, 44

2.3

Developments leading to Bohr's Model

2.3 Developments leading to Bohr's Model

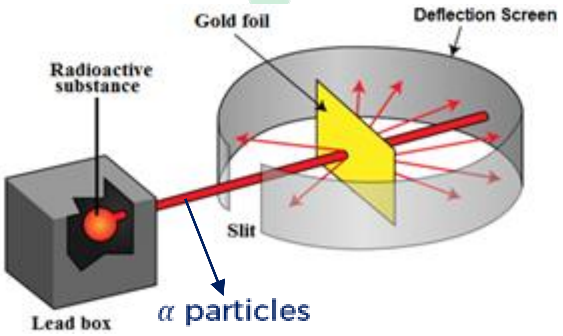
Learning Objectives

- Electromagnetic Wave
- Plank quantum theory
- Photoelectric effect
- Atomic spectra

2.3.1 Electromagnetic Wave

Drawbacks of Rutherford Model

- Stability of Electron ✖
- Atomic Spectra ✖
- Distribution of electrons ✖

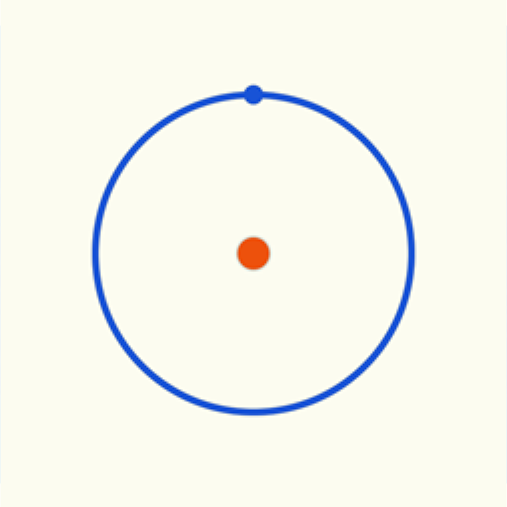


Rutherford's Model

Drawbacks of Rutherford Model

1. Stability of Electron ✖

- Electrons are charged whereas planets are uncharged
- According to the electromagnetic theory of Maxwell, charged particles when accelerated should emit electromagnetic radiation
- The orbit will thus continue to shrink
- Calculations show that it should take an electron only $10^{-8}\,s$ to spiral into the nucleus.

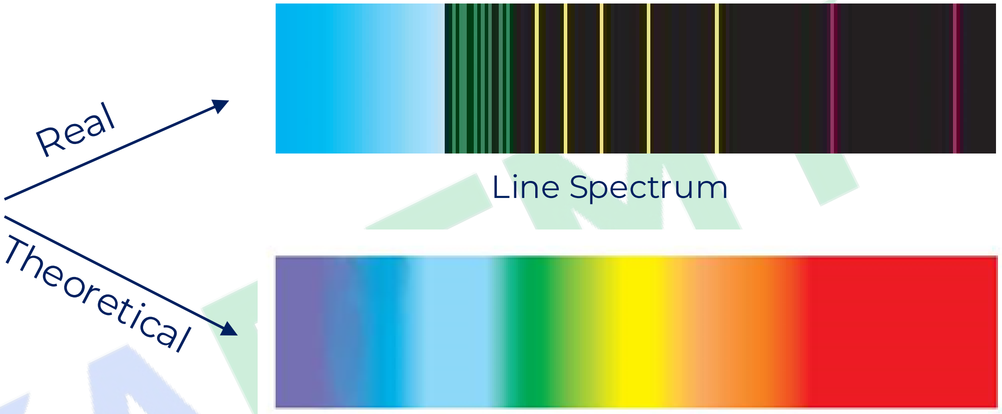


Planetary model
nucleus ≡ Sun
electrons ≡ Planets

Drawbacks of Rutherford Model

1. Stability of Electron

2. Atomic Spectra ✖



Continuous Spectrum

Acc. to electromagnetic theory

Accelerated electron → Continuously emits energy

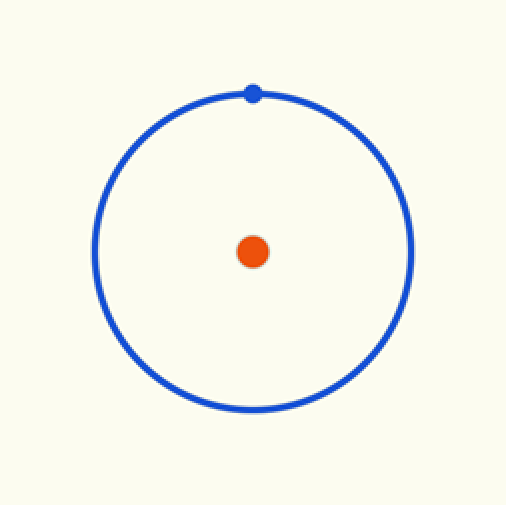
Drawbacks of Rutherford Model

1. Stability of Electron

2. Atomic Spectra

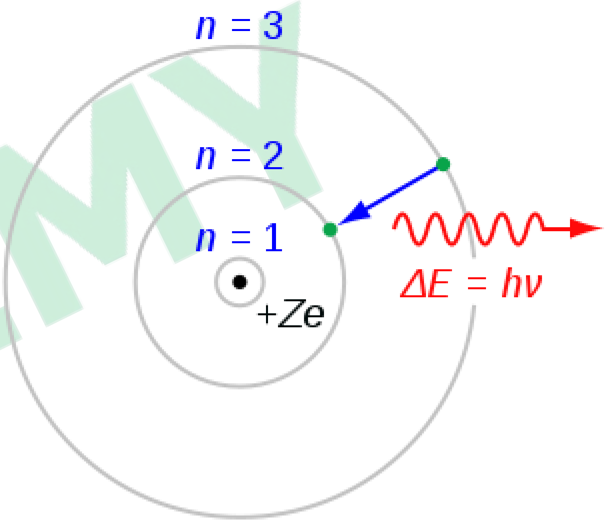
3. Distribution of electrons ✖

Distribution of electron



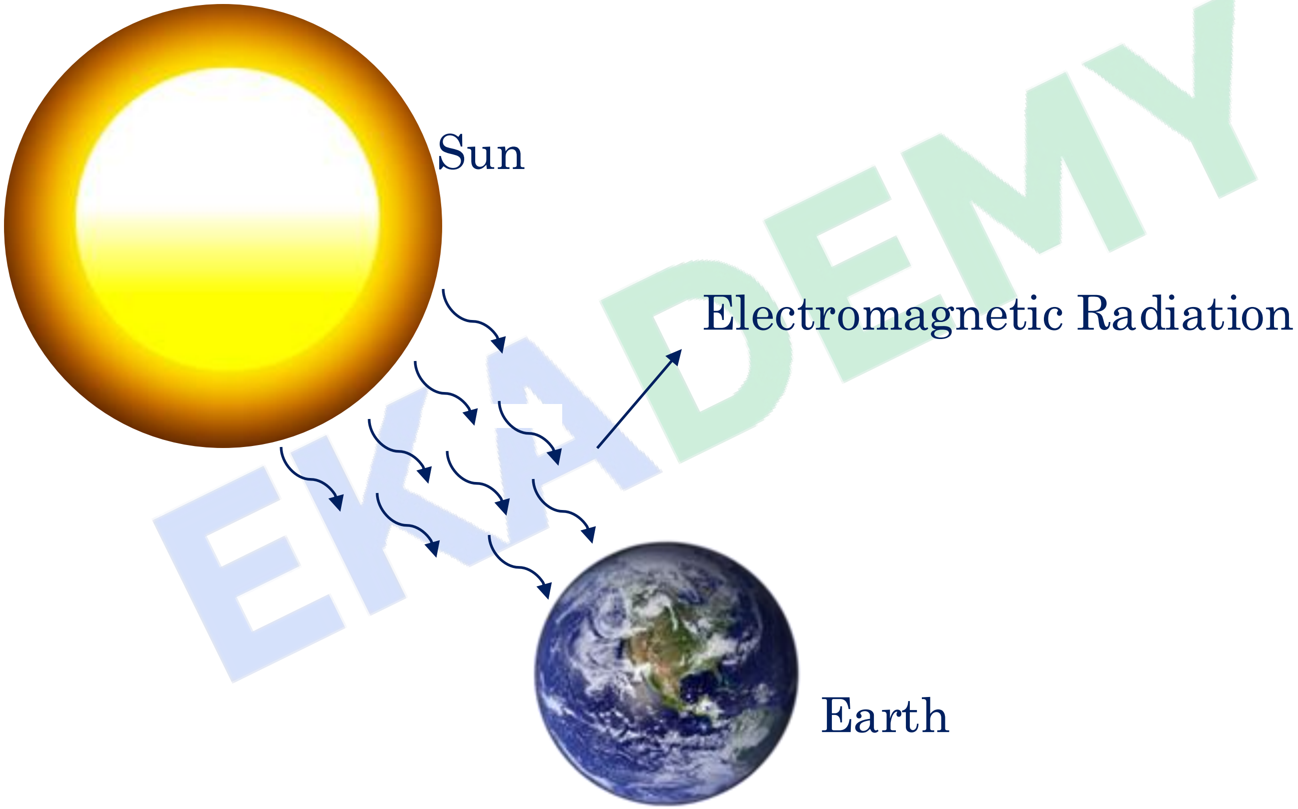
Rutherford's Model

Next →



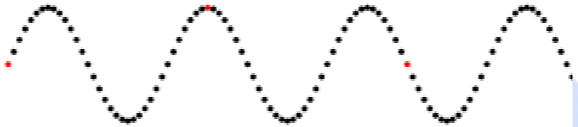
Bohr's Model

Electromagnetic Radiation :

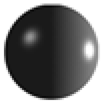


Nature of Electromagnetic Radiation

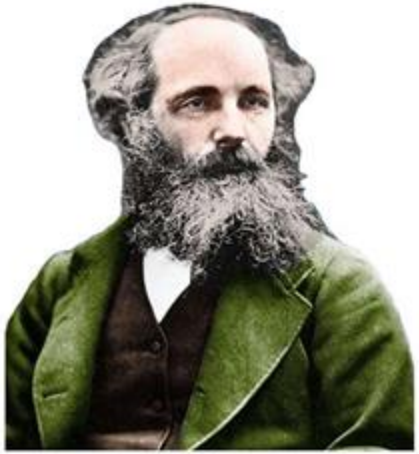
Wave Nature



Particle Nature



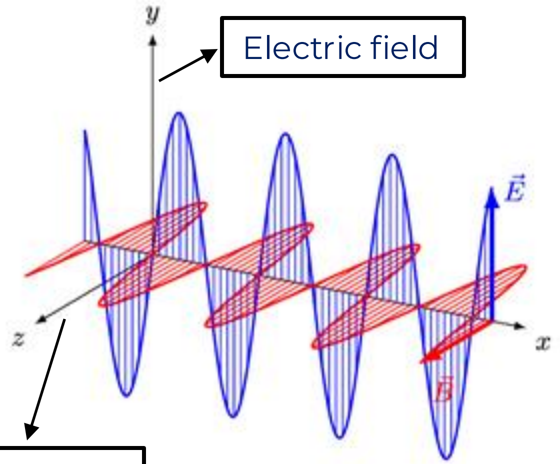
Electromagnetic Radiation



James Maxwell

Wave Nature

Electric field



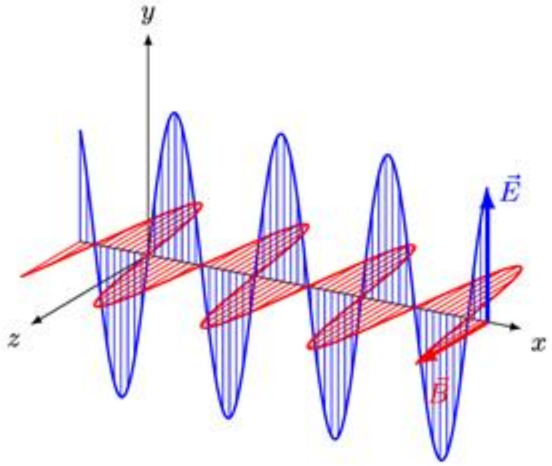
Magnetic field

Properties of electromagnetic Waves

1. Oscillating Magnetic field $\perp$ Electric field

   Propagation of waves $\perp$ Magnetic field & Electric field

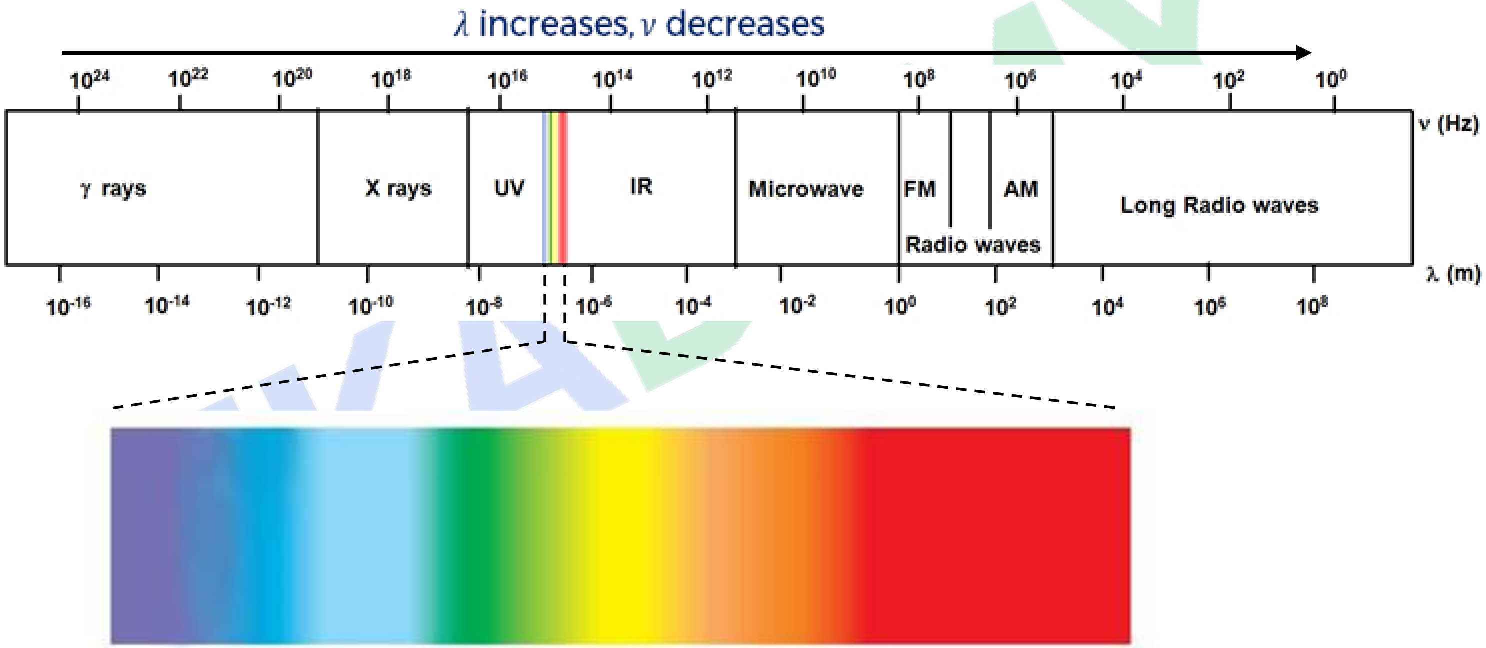
2. EM Waves can travel in vacuum



3. Speed $\longrightarrow$ Light $(3 \times 10^8 \frac{\text{m}}{\text{s}})$

Properties of electromagnetic Waves

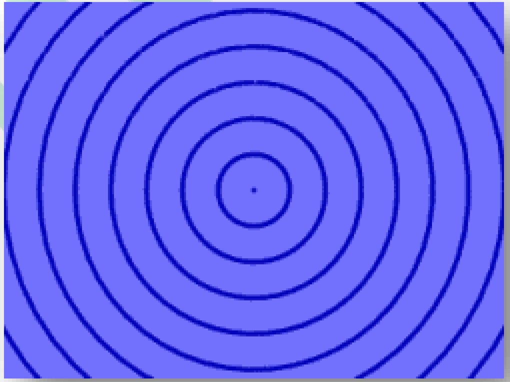
4 . Electromagnetic spectrum



Visible spectrum

Representation of Electromagnetic waves

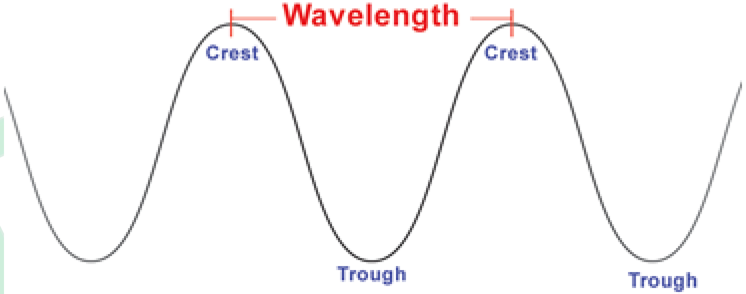
- Wavelength (λ)
- Frequency (ν)
- Wave Number ($\overline{\nu}$)



Water waves

Representation of Electromagnetic waves

- Wavelength (λ)
- Frequency (ν)
- Wave Number ($\overline{\nu}$)

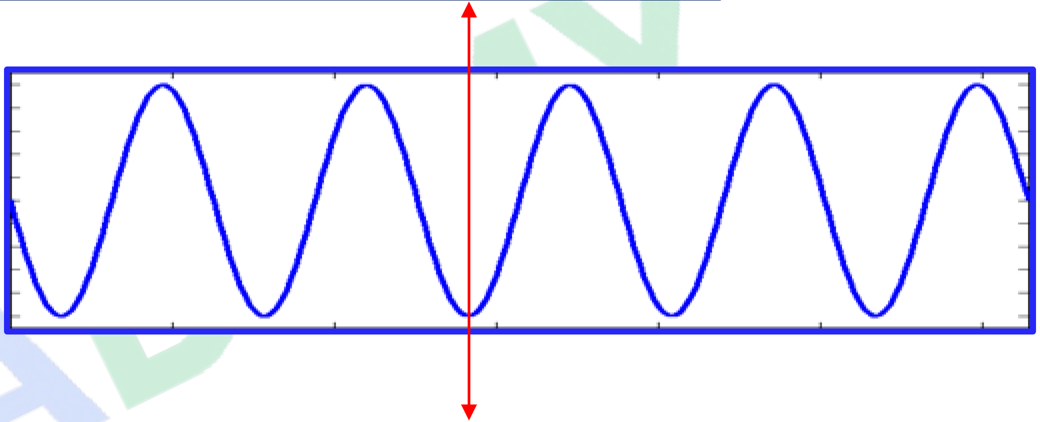


λ = distance b/w two consecutive Crest or Trough

SI unit = m

Representation of Electromagnetic waves

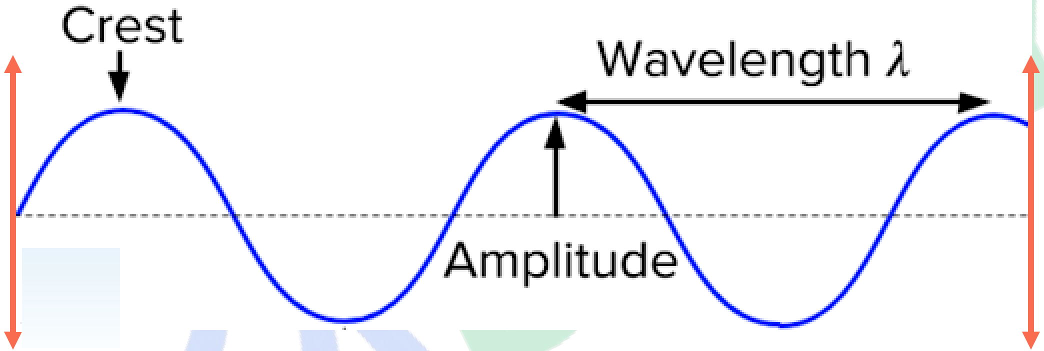
- Wavelength (λ)
- Frequency (ν)
- Wave Number ($\bar{\nu}$)



ν = No. of waves / sec

SI Unit = Sec^{-1} or Hertz

Relation between speed , frequency and wavelength

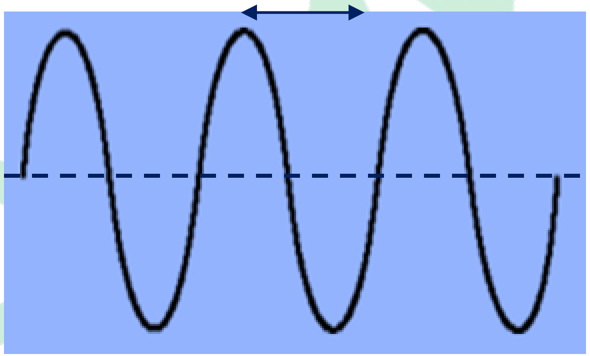


Speed = Frequency × Wavelength

$$c = \nu \times \lambda$$

Representation of Electromagnetic waves

- Wavelength (λ)
- Frequency (ν)
- Wave Number ($\bar{\nu}$)



$\bar{\nu}$ = No. of wavelength / unit length

$$\bar{\nu} = \frac{1}{\lambda}$$

SI Unit = m^{-1}

Solved Problems

Q. Calculate the wave number and frequency of violet radiation having wavelength 4000 Å.

Sol:

Given : $\lambda = 4000\ \text{Å}$

$\lambda = 4 \times 10^{-7}\ \text{m}$

$\boxed{1\text{Å} = 10^{-10}\ \text{m}}$

$\boxed{\bar{\nu} = 1/\lambda}$

$$\bar{\nu} = \frac{1}{4 \times 10^{-7}}\ \text{m}^{-1}$$

$\bar{\nu} = 2.5 \times 10^{6}\ \text{m}^{-1}$

$\boxed{c = \nu\lambda}$

$\boxed{c = 3 \times 10^{8}\ \frac{\text{m}}{\text{s}}}$

$3 \times 10^{8} = \nu \times 4 \times 10^{-7}$

$\nu = 7.5 \times 10^{14}\ \text{sec}^{-1}$

Concept Test

Ready for the challenge

Q. Calculate the wavelength, frequency of a light wave whose period is 2.0×10^{-10} s.

Pause the video

Time duration: 2 minutes

Q. Calculate the wavelength, frequency of a light wave whose period is 2.0×10^{-10} s.

Sol: Given : $T = 2.0 \times 10^{-10}$ s

$$\boxed{\nu = \frac{1}{T}}$$

$$\nu = \frac{1}{2\times10^{-10}} \sec^{-1}$$

$$\nu = 5 \times 10^{9} \sec^{-1}$$

$$\boxed{\lambda = \frac{c}{\nu}}$$

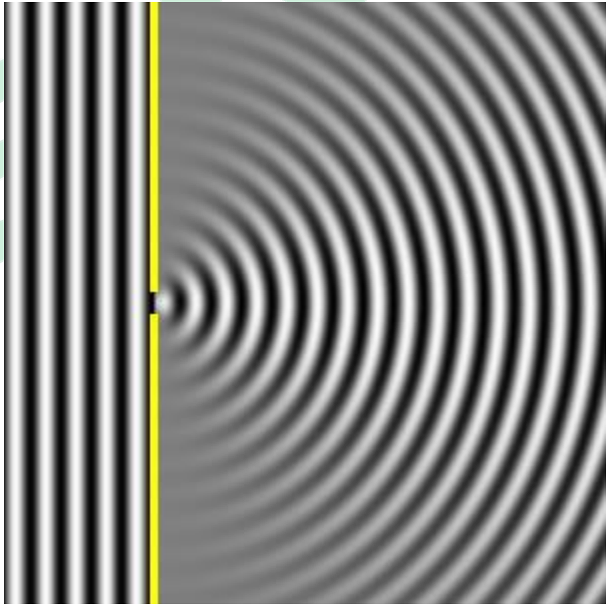
$$\lambda = \frac{3 \times 10^{8}}{5 \times 10^{9}} \text{ m}$$

$$\lambda = 6 \times 10^{-2} \text{ m}$$

2.3.2
Planck's Quantum Theory

Wave nature of Electromagnetic radiation

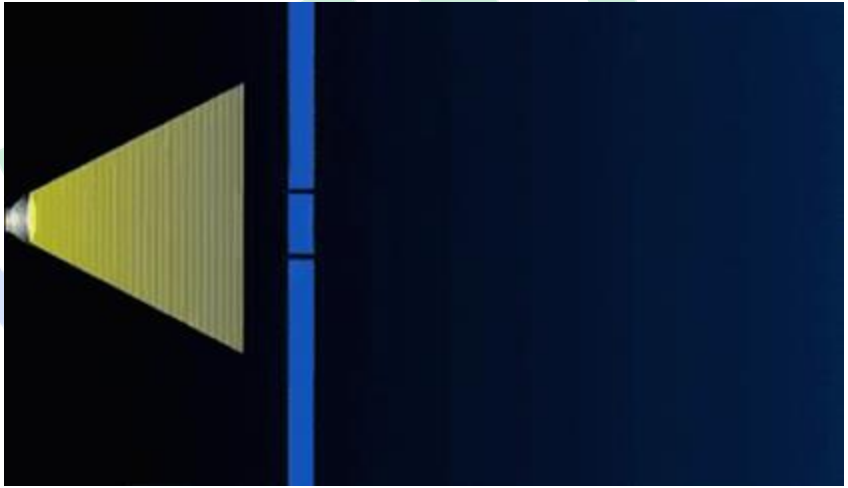
Diffraction ✓



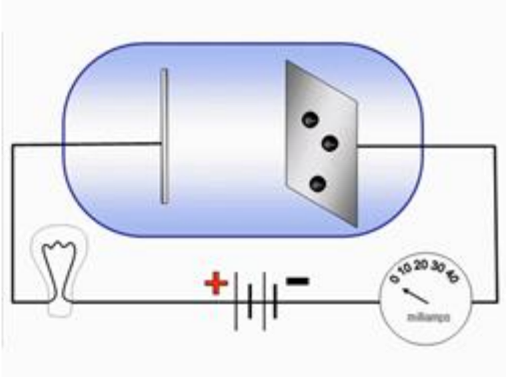
Wave nature of Electromagnetic radiation

Diffraction ✔

Interference ✔



Wave nature of Electromagnetic radiation



- ✖ Black body radiation
- ✖ Photoelectric effect
- ✖ Line spectrum (H)
- ✖ Variation of Heat capacity of solids as $f\ (T)$

Black body radiation

Black Shirt

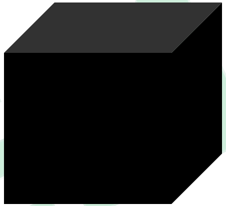
Comfortable in Hot

Comfortable in Hot

White Shirt

Black body radiation

Black body

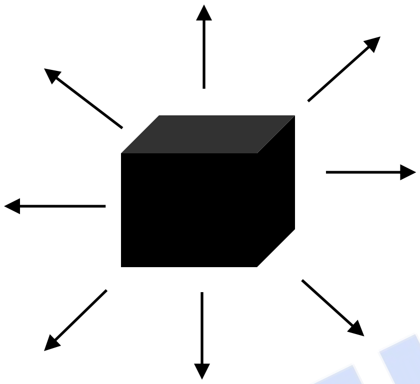


Emits and Absorbs → **All kinds Frequency uniformly**

Practically → 

Black body ≈ Carbon black

Black body radiation

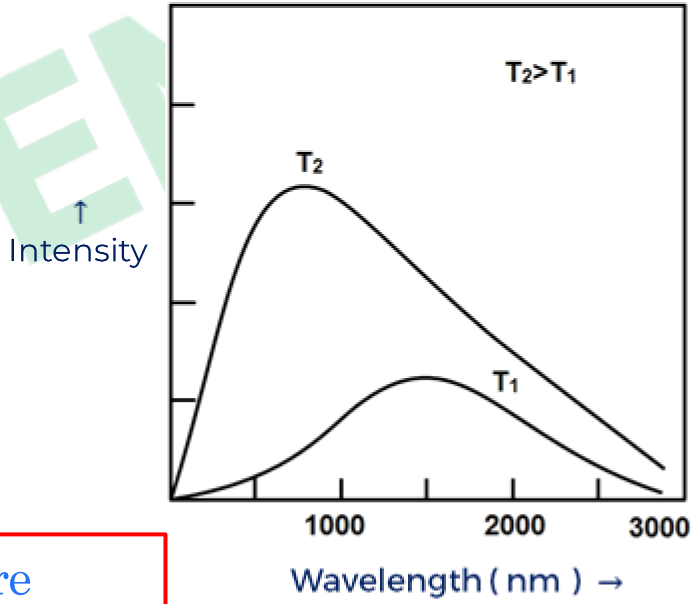


Black body

Intensity depends :

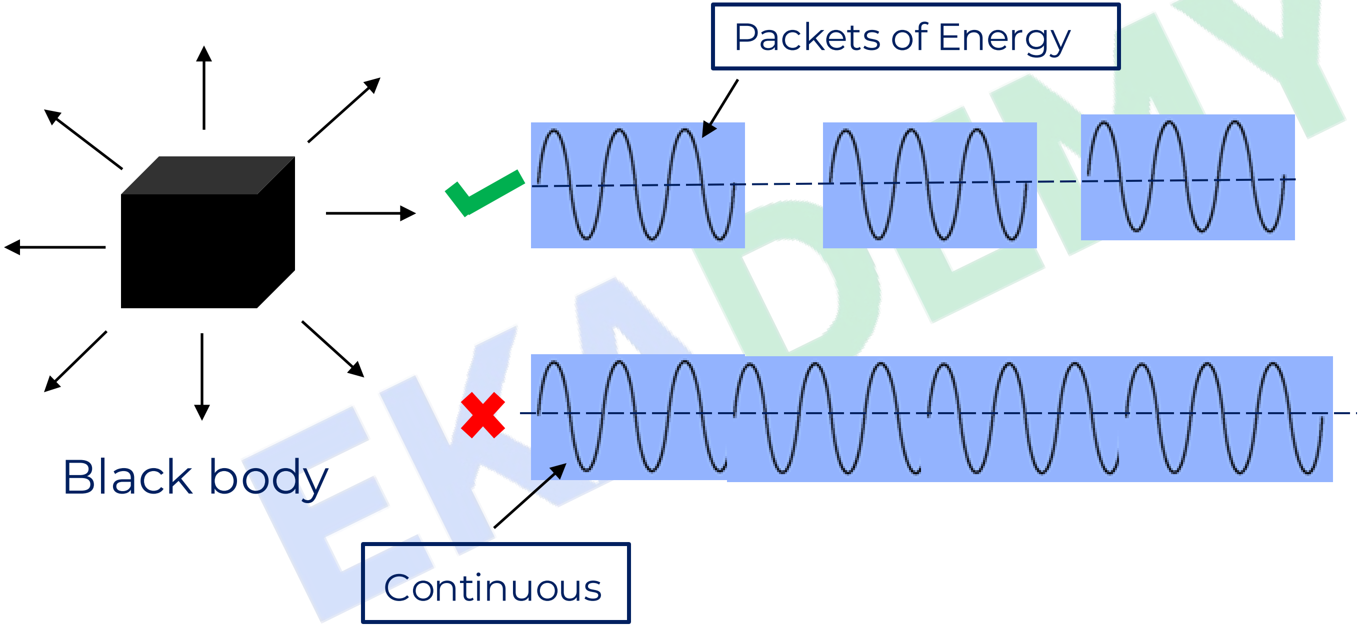
→ Temperature

→ Wavelength



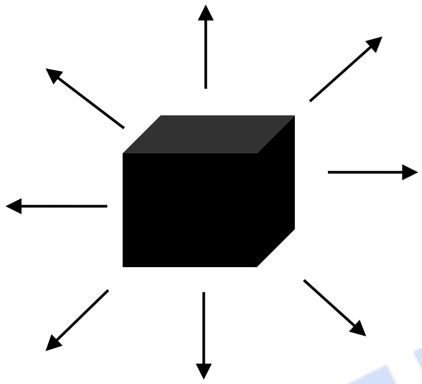
Unexplained by Wave nature

Planck 's quantum Theory



Max. Planck

Planck 's quantum Theory



Black body

Packet of Energy $\longrightarrow$ Quantum

Energy $\propto$ Frequency

$$E \propto \nu$$

$$\boxed{E = h\nu}$$

h = Planck 's constant $= 6.626 \times 10^{-34}$ J s

$$E = 0, h\nu, 2h\nu, 3h\nu \ldots nh\nu \ldots$$

Max. Planck

Solved Problems

Q. 100 Watt bulb emits the monochromatic light of 400nm . Calculate the number of photons emitted per second by the bulb .

Sol: Given :

$$\lambda = 400 \text{ nm}$$

$$\lambda = 4 \times 10^{-7} \text{ m} \qquad 1 \text{ nm} = 10^{-9} \text{ m}$$

100 watt = 100 j / sec

$$\boxed{100 \text{ j / sec} = \text{Energy of n number photons / sec}}$$

$$100 = n h\nu$$

$$100 = n\frac{hc}{\lambda}$$

$$n = \frac{100\lambda}{hc}$$

$$n = \frac{100 \times 4 \times 10^{-7}}{6.626 \times 10^{-34} \times 3 \times 10^{8}}$$

$$\boxed{n = 2.012 \times 10^{20} \text{ photons}}$$

$$c = 3 \times 10^{8} \frac{\text{m}}{\text{s}}$$

$$h = 6.626 \times 10^{-34} \text{ J s}$$

Concept Test

Ready for challenge

Q. Calculate the energy of 1 mole photons of radiation having frequency 5×10^{14} Hz .

Pause the video

Time duration: 2 minutes

Q. Calculate the energy of 1 mole photons of radiation having frequency 5×10^{14} Hz .

Sol: Given : $\nu = 5 \times 10^{14}\,\text{sec}^{-1}$

$h = 6.626 \times 10^{-34}$ J s

According to Planck's quantum theory :

Energy of 1 photon = $h\nu$

Energy of 1 mole photons = $N_A h\nu$

$N_A = 6.023 \times 10^{23}$

$E = 6.023 \times 10^{23} \times 6.626 \times 10^{-34} \times 5 \times 10^{14}$

$E = 199.5$ KJ/mole

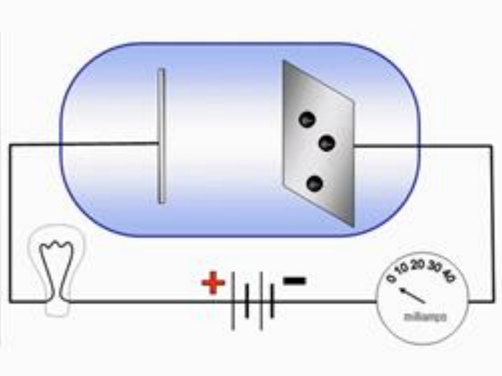
2.3.3
Photoelectric Effect

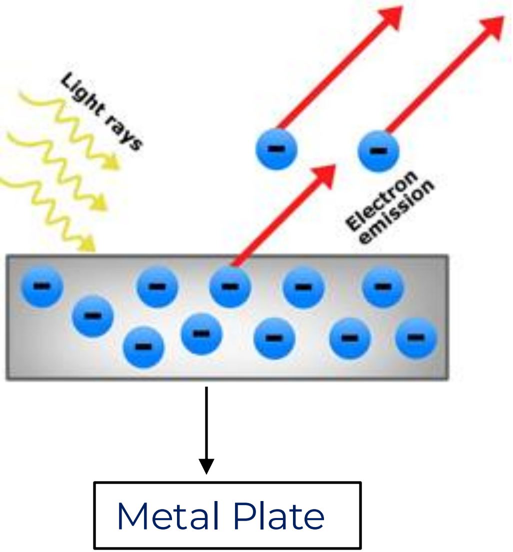
Photoelectric Effect

H. Hertz

For instance :

Potassium , Rubidium etc





Photoelectric Effect

Result observed :

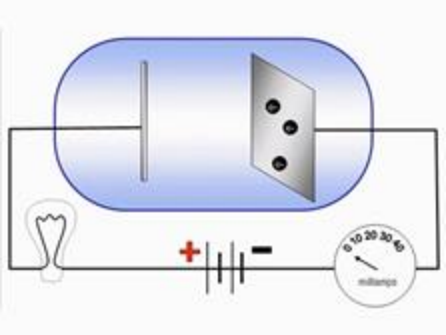
Brightness

No. of Electron ejected ∝ $\boxed{\text{Intensity of Light}}$

No time lag b/w light striking & Electron ejection

Electron ejection (Metals) → $\boxed{\text{Minimum frequency}}$

Photoelectric effect ✖ ⟵ Frequency < **Threshold Frequency**



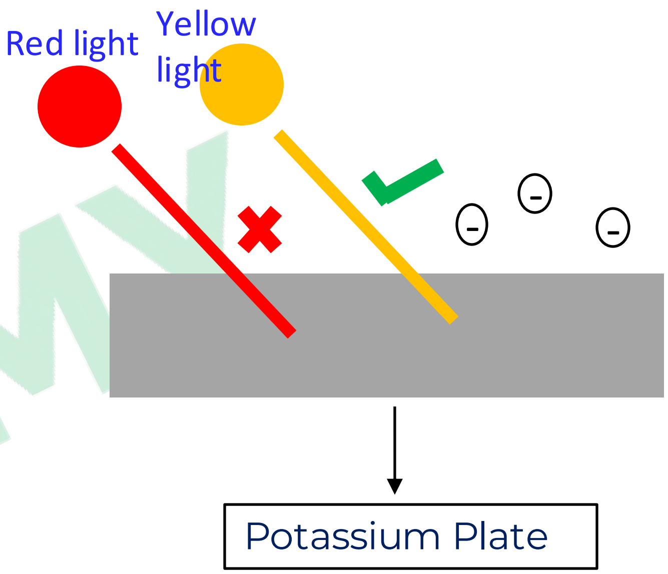
Observations could not be explained by wave nature of light

Photoelectric Effect

No. of Electron ejected depends → Intensity of light

Energy of electron ~~→ Intensity of light~~

Energy of electron → Frequency of light



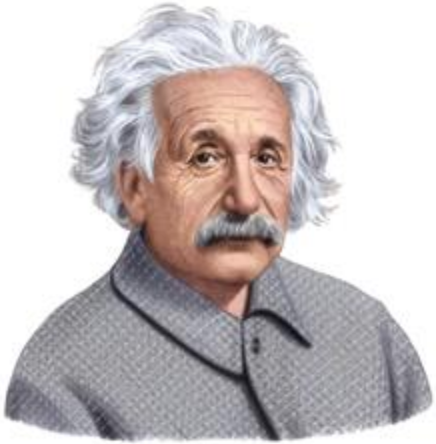
If light is a wave or continuous form of energy than:

- there should always be emission of electron, no matter the frequency
- even if the intensity of light is low, with time there will be ejection of electron
- there will be no threshold frequency

Photoelectric Effect

↓

Quantum Theory of Electromagnetic Radiation



Albert Einstein

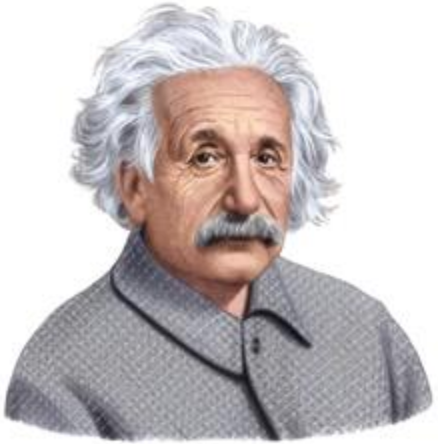
Max. Planck

Threshold frequency → ν_o

Minimum energy to eject electron → $h\nu_o$

Work function

Photoelectric Effect



Albert Einstein

Photon frequency → ν

Photon energy → $h\nu$

$h\nu - h\nu_o$ = Kinetic energy of electron

$$\boxed{h\nu - h\nu_o = \frac{1}{2}m_e v^2}$$

Velocity of ejected electron

Mass of electron

Dual Behaviour of Electromagnetic Radiation

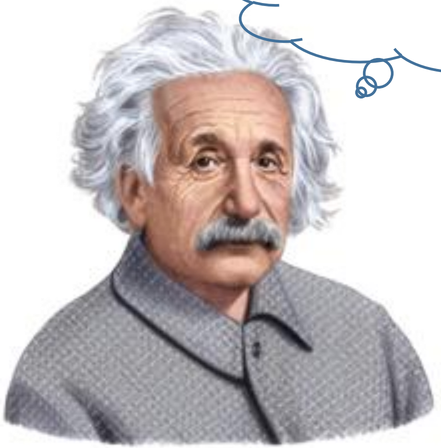
Photoelectric Effect

Interference and diffraction

Particle nature

Dilemma??

Wave nature



Albert Einstein

Dual nature of light

Solved Problems

Q. The threshold frequency ν_o for a metal $6 \times 10^{14}\ s^{-1}$. Calculate the kinetic energy of emitted electron when radiation of frequency $1.1 \times 10^{15}\ s^{-1}$.

Sol:

Given :

$$\nu_o = 6 \times 10^{14}\ s^{-1}$$

$$\nu = 1.1 \times 10^{15}\ s^{-1}$$

$$\boxed{h\nu - h\nu_o = K.E.}$$

$$h(\nu - \nu_o) = K.E. \qquad h = 6.626 \times 10^{-34}\ J\,s$$

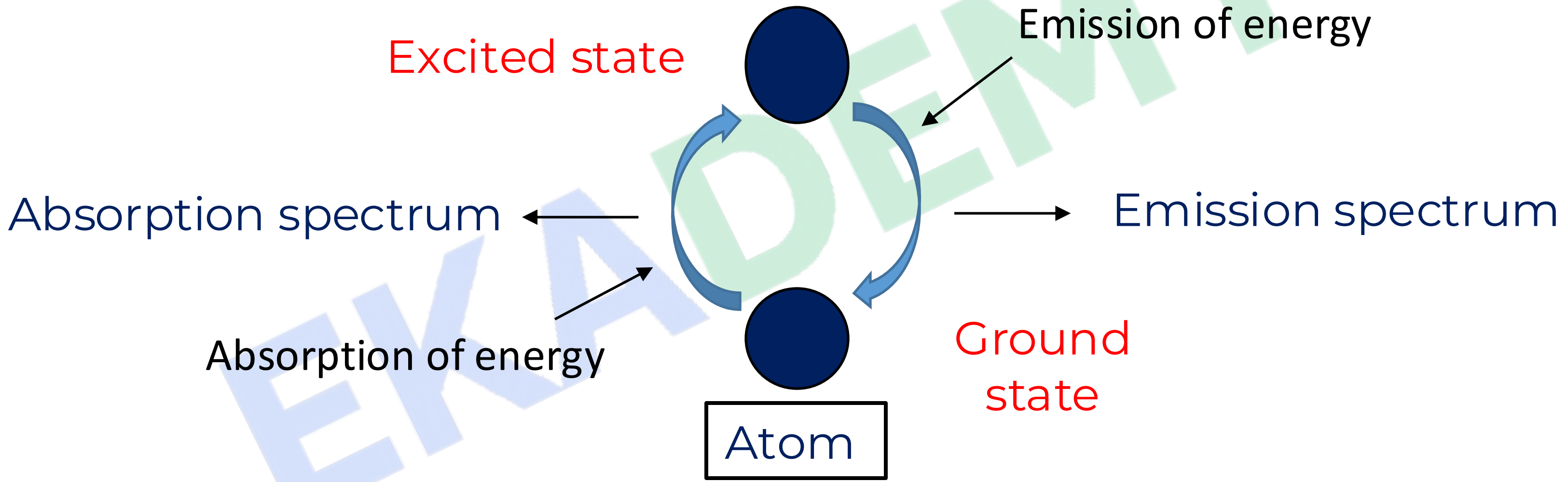
$$6.626 \times 10^{-34}(1.1 \times 10^{15} - 6 \times 10^{14}) = K.E.$$

$$K.E. = 33.13 \times 10^{20}\ J$$

2.3.4
Line spectrum of Hydrogen

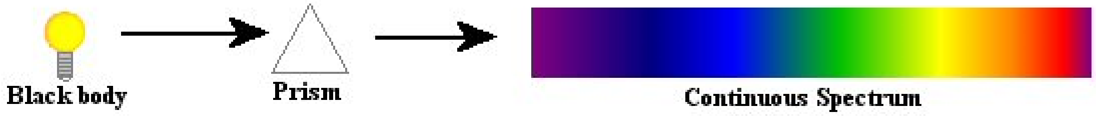
Spectrum

Spectrum


Excited state
Emission of energy
Absorption spectrum
Emission spectrum
Absorption of energy
Ground state
Atom

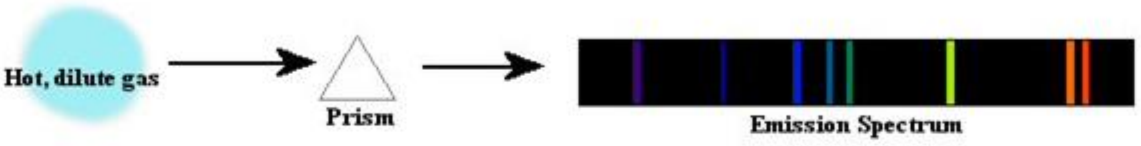
Spectrum

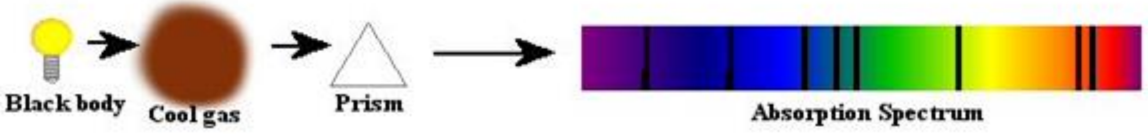
- Black body Spectrum



Continuous Spectrum

- Gaseous atom Spectrum





Line Spectrum

Line Spectrum of Hydrogen

Balmer

Hydrogen spectrum → Several series of Line → Wave Number ($\overline{\upsilon}$)

$$\overline{\nu} = 109{,}677 \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \text{cm}^{-1}$$

Balmer formula

109,677 → Rydberg Constant

n → n = 3, 4, 5........

Line Spectrum of Hydrogen

Hydrogen spectrum → All series of Line →

- Lyman Series
- Balmer Series
- Paschen Series
- Brackett Series
- Pfund Series

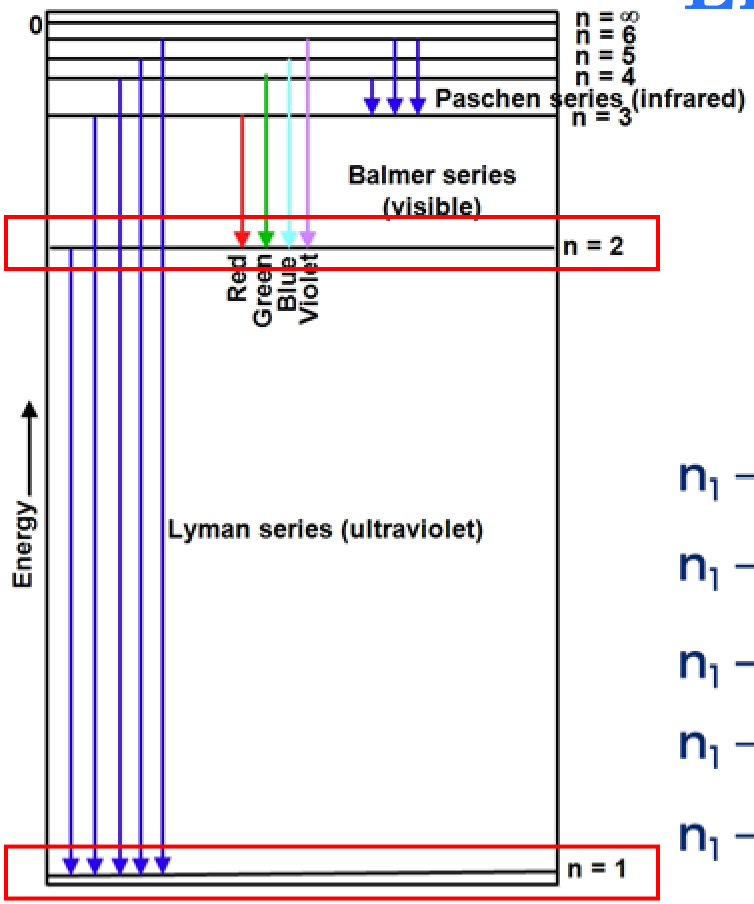
$$\bar{\nu} = 109{,}677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)\ \mathrm{cm^{-1}}$$

Rydberg formula

$n_1 = 1, 2 \ldots\ldots\ldots$

$n_2 = n_1+1, n_1+2 \ldots\ldots\ldots$

Line Spectrum of Hydrogen



$$\boxed{\bar{\nu} = 109{,}677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)\ \text{cm}^{-1}}$$

$n_1 \rightarrow 1 \rightarrow$ Lyman series $\rightarrow$ Ultraviolet

$n_1 \rightarrow 2 \rightarrow$ Balmer series $\rightarrow$ Visible

$n_1 \rightarrow 3 \rightarrow$ Paschen series $\rightarrow$ Infrared

$n_1 \rightarrow 4 \rightarrow$ Brackett series $\rightarrow$ Infrared

$n_1 \rightarrow 5 \rightarrow$ Pfund series $\rightarrow$ Infrared

$\boxed{n_1 < n_2}$

Line Spectrum of Hydrogen

Series	n_1	n_2	Spectral Region
Lyman	1	2,3 ...	Ultraviolet
Balmer	2	3,4 ...	Visible
Paschen	3	4,5 ...	Infrared
Brackett	4	5,6 ...	Infrared
Pfund	5	6,7 ...	Infrared

Solved Problems

Q. What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from an energy level with n = 4 to an energy level with n = 2?

Sol:

Given : $n_1 = 2\ ,n_2 = 4$

$\lambda = \ ?$

Rydberg formula → $\boxed{\bar{\nu} = 109{,}677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)\ \mathrm{cm^{-1}}}$

$$\bar{\nu} = 109{,}677\left(\frac{1}{2^2} - \frac{1}{4^2}\right)\ \mathrm{cm^{-1}}$$

$$\bar{\nu} = 109{,}677\left(\frac{1}{4} - \frac{1}{16}\right)\ \mathrm{cm^{-1}}$$

$$\bar{\nu} = 20564.43\ \mathrm{cm^{-1}}$$

$$\lambda = \ 4.86\ \times 10^{-5}\ \mathrm{cm}$$

$$\boxed{\bar{\nu} = \frac{1}{\lambda}}$$

$$\lambda = \ 4.86\ \times 10^{-7}\ \mathrm{m}$$

$$\boxed{\lambda = \ 486\ \mathrm{nm}}$$

Summary

- Electromagnetic Waves : Characteristics of waves is wavelength , frequency and wave number

- Planck quantum theory : $E = h\nu$

- Photoelectric effect : $h\nu - h\nu_o = \frac{1}{2} m_e v^2$

- Atomic Spectra : $\overline{\nu} = 109{,}677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{cm}^{-1}$

2.3 Developments leading to Bohr's Model

Reference Questions

NCERT Exercises: 5, 6, 7, 8, 9, 11, 12, 13, 17, 33, 51, 52, 53, 54

2.4

Bohr's Model for Hydrogen Atom

2.4 Bohr's Model for Hydrogen Atom

Learning Objectives

- Postulates of Bohr's Model
- Results obtained from Bohr's Model
- Explanation of line spectrum of Hydrogen
- Limitations of Bohr's Model

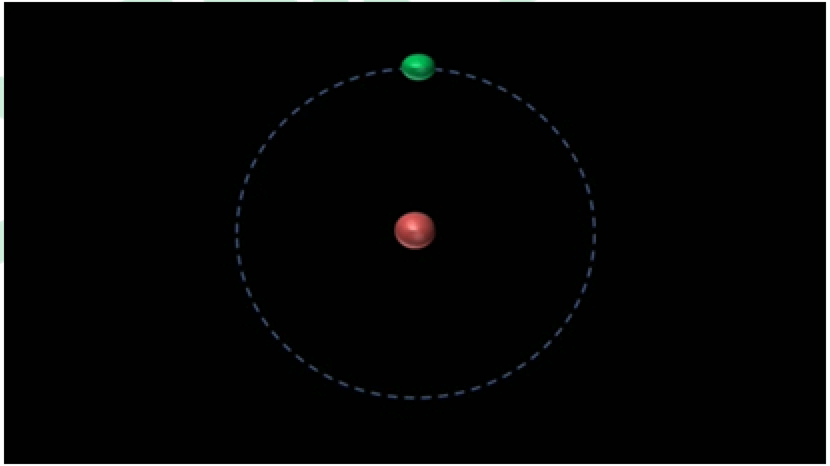
2.4.1

Postulates of Bohr's Model

Introduction to Bohr's Model

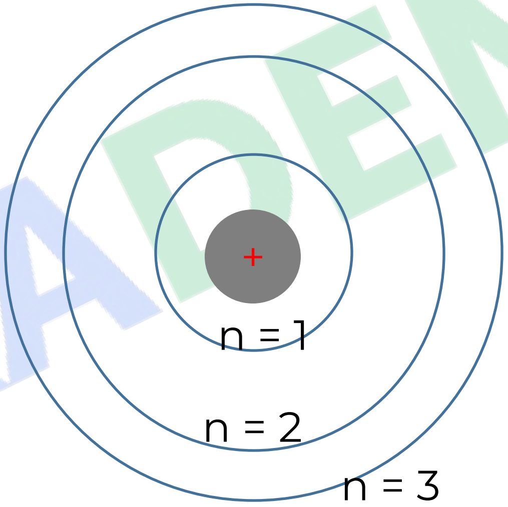
General features of the structure of hydrogen atom and its spectrum is explained

Neils Bohr



Postulates of Bohr's Model

1. Electron revolve around nucleus in fixed circular paths(Orbits).
   These Orbits have constant energy and radius



It removed the limitation of collapsing of electron in Rutherford Atomic Model

Postulates of Bohr's Model

2. The orbit of electron does not change unless energy is added or removed

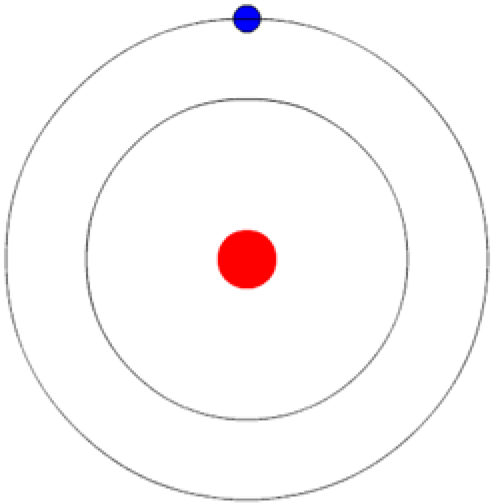
Gain energy → Jumps from lower to higher n

Lose/ emit energy → Jumps from higher to lower n

Postulates of Bohr's Model

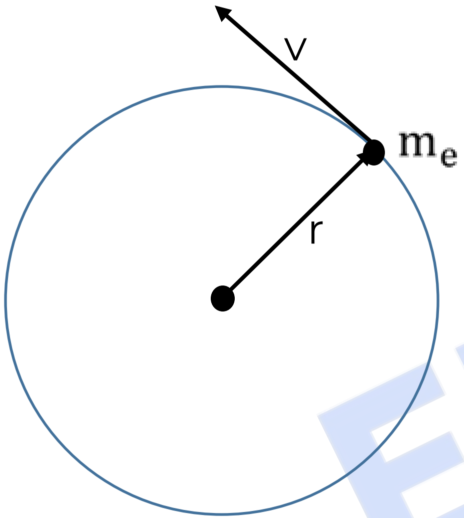
3. The frequency of the radiation emitted/absorbed by an atom depends upon the energy difference of the initial & final orbits

$$\nu = \frac{\Delta E}{h} = \frac{E_{final} - E_{initial}}{h}$$



Postulates of Bohr's Model

4. The angular momentum of an electron is quantized



$$m_e v r = n.\frac{h}{2\pi}$$

n = Integer

Electron can move only in those orbits for which its angular momentum is integral multiple of $\frac{h}{2\pi}$

Concept Test

Ready for Challenge

Q. Which of the following cannot be the value of angular momentum of electron in Hydrogen atom ?

A. $\frac{h}{2\pi}$

B. $\frac{h}{\pi}$

C. $\frac{3h}{4\pi}$

D. $\frac{3h}{2\pi}$

Pause the video

Time duration : 1 minute

Q. Which of the following cannot be the value of angular momentum of electron in Hydrogen atom ?

Sol.

A. $\frac{h}{2\pi} = 1.\frac{h}{2\pi}$

B. $\frac{h}{\pi} = 2.\frac{h}{2\pi}$

C. $\frac{3h}{4\pi} = \frac{3}{2}.\frac{h}{2\pi}$

D. $\frac{3h}{2\pi} = 3.\frac{h}{2\pi}$

Representing in terms of $n.\frac{h}{2\pi}$

$\cancel{n = \frac{3}{2}}$

Not an Integer

Ans. C

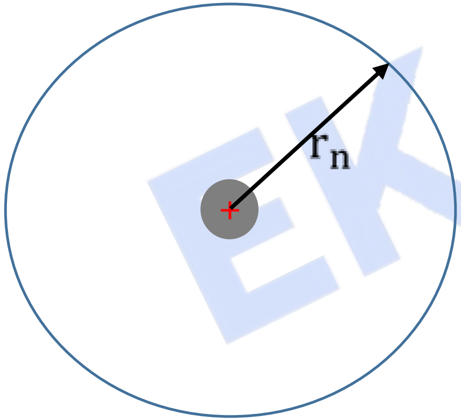
2.4.2

Results obtained from Bohr's Model

Results obtained from Bohr's model

1. The stationary states(orbits) for electron are numbered n = 1,2,3……….
   These integral numbers are known as Principal quantum numbers

2. The radii of the stationary states are expressed as :



$$r_n = n^2 a_0 \qquad where\ a_0 = 52.9\ pm$$

3. The energy associated with electron in its stationary state/orbit/shell is given by :

$$E_1 = -2.18 \times 10^{-18}\left(\frac{1}{1^2}\right) J = -2.18 \times 10^{-18} \ J$$

$$E_n = -R_H\left(\frac{1}{n^2}\right)$$

$$E_2 = -2.18 \times 10^{-18}\left(\frac{1}{2^2}\right) J = -0.545 \times 10^{-18} \ J$$

R_H = Rydberg constant

$= 2.18 \times 10^{-18}$ J

.

.

.

.

$$E_\infty = -2.18 \times 10^{-18}\left(\frac{1}{\infty^2}\right) J = 0 = \text{energy of free electron at rest}$$

Note: Energy associated with free electron at rest is zero

4. Bohr's theory is applicable to the ions containing only one electron. For example, $He^+, Li^{2+}, Be^{3+}, etc.$

Energies & radii of the stationary states associated with these kinds of ions are given by :

Energy

$$E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2}\right) J$$

Z = Atomic Number

Radius

$$r_n = 52.9 \left(\frac{n^2}{Z}\right) pm$$

5. Velocities of an electron moving in a stationary orbit can be calculated as :

$$V_n = 2.165 \times 10^6 \left(\frac{Z}{n}\right) \text{ m s}^{-1}$$

$$V_n \propto \frac{1}{n}$$

$$V_n \propto Z$$

Solved Problems

Q. What is the radius of third orbit of the hydrogen atom ?

Q. What is the radius of third orbit of the hydrogen atom ?

Sol.

We know that :

Radius of orbit of Hydrogen is given by :

$$r_n = 52.9\left(\frac{n^2}{Z}\right)\text{pm}$$

Here, n = 3, Z = 1

$$\Rightarrow \quad r_n = 52.9\left(\frac{3^2}{1}\right)\text{pm}$$

$$= 476.1\ \text{pm}$$

Q. Calculate the difference in energy of electron, between 4^{th} and 2^{nd} orbit of hydrogen atom.

Q. Calculate the difference in energy of electron, between 4^{th} and 2^{nd} orbit of hydrogen atom.

Sol.

$$\Delta E = E_4 - E_2$$

$$= \left(-\frac{R_H}{n_4^2}\right) - \left(-\frac{R_H}{n_2^2}\right)$$

$$= R_H\left(\frac{1}{n_2^2} - \frac{1}{n_4^2}\right)$$

$$= 2.18 \times 10^{-18}\left(\frac{1}{2^2} - \frac{1}{4^2}\right) J$$

$$= 0.40875 \times 10^{-18} J$$

2.4.3

Explanation of Line Spectrum of Hydrogen

Explanation of Line Spectrum of Hydrogen

Each spectral line, whether in absorption/ emission spectrum, can be associated to the particular transition in hydrogen atom

Energy

$$\Delta E = E_f - E_i = \left(-\frac{R_H}{n_f^2}\right) - \left(-\frac{R_H}{n_i^2}\right) = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\Delta E = 2.18 \times 10^{-18}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \text{ J}$$

where n_i & n_f stands for initial orbit and final orbits

This expression is similar with Rydberg expression

Explanation of Line Spectrum of Hydrogen

Frequency(ν) :

$$\nu = \frac{\Delta E}{h} = \frac{R_H}{h}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

Wavenumber($\bar{\nu}$) :

$$\bar{\nu} = \frac{\nu}{c} = \frac{R_H}{hc}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

In case of large number of hydrogen atoms, different possible transitions can be observed and thus leading to large number of spectral lines

The intensity of spectral lines depends upon the number of photons of same wavelength or frequency absorbed/ emitted

2.4.4

Limitations of Bohr's Model

Limitations of Bohr's Model

Bohr's model was a big improvement over Rutherford's nuclear model,
as it could account for :

- Stability
- Line spectra

of hydrogen atom and hydrogen like ions

But...

Limitations of Bohr's Model

Bohr's model was unable to explain the following points :

- Details of the hydrogen atom spectrum.
- Spectrum of atoms having more than one electron.
- Splitting of spectral lines in the presence of magnetic field (Zeeman effect) or an electric field (Stark effect).
- Ability of atoms to form molecules by chemical bonds.

Summary

Postulates of Bohr's Model : Angular momentum is quantized

Results obtained from Bohr's Model :

$$r_n = n^2 a_o$$

$$E_n = -R_H\left(\frac{1}{n^2}\right)$$

$$V_n = 2.165 \times 10^6 \left(\frac{Z}{n}\right) \text{ m s}^{-1}$$

Explanation of line spectrum of Hydrogen

Limitations of Bohr's Model

2.4 Reference questions

NCERT Exercise questions : 2.13, 2.14, 2.16, 2.18, 2.19, 2.33

2.5

Towards Quantum Mechanical Model of the Atom

2.5 Towards Quantum Mechanical Model of the Atom

Learning Objectives

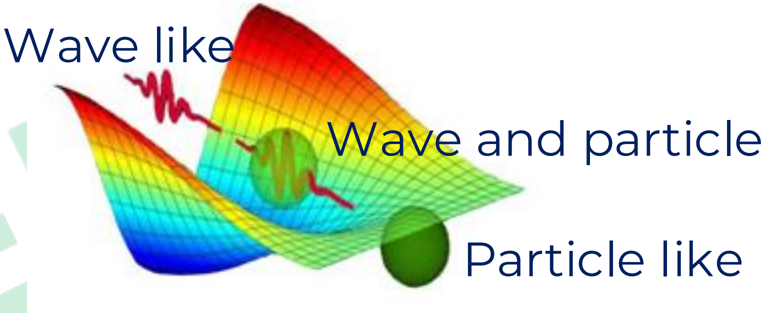
- Dual Behaviour of Matter
- Heisenberg's Uncertainty Principle

2.5.1

Dual Behaviour of Matter

Dual Behaviour of Matter

Louis de Broglie



Matter → Particle properties

Matter → Wavelike properties

Dual Behaviour of Matter

Louis de Broglie

Particle nature ← Electron beam → Diffraction

↓

Wave nature

↓

Develop electron microscope

Dual Behaviour of Matter

Louis de Broglie

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

h = Plank's constant

λ = Wavelength

p = Momentum of a material particle.

Wavelength

Short

For ordinary object like ball etc.

Because large masses

Can be detected experimentally.

For electrons and other subatomic particles

Because very small masses

Solved Problem

Q. Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.

Q. Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.

Sol

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

$$h = 6.626 \times 10^{-34}\ \text{kg m}^2\text{s}^{-1}$$

$$\lambda = 8 \times 10^{-10}\ \text{m}$$

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

$$m = 1.675 \times 10^{-27}\ \text{kg}$$

$$v = \frac{6.626 \times 10^{-34}}{1.675 \times 10^{-27} \times 8 \times 10^{-10}}$$

$$v = 494\ \text{m s}^{-1}$$

Concept Test

Q. Find the wavelength of a particle of 100 g moving with velocity 100 m s^{-1}

Q. Find the wavelength of a particle of 100 g moving with velocity 100 m s^{-1}

Sol.

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

$$h = 6.626 \times 10^{-34}\ kg\ m^2 s^{-1}$$

$$m = 0.1\ Kg$$

$$v = 100\ m\ s^{-1}$$

$$\lambda = \frac{6.626 \times 10^{-34}}{0.1 \times 100}$$

$$\lambda = 6.626 \times 10^{-35}\ m$$

2.5.2

Heisenberg's Uncertainty Principle

Heisenberg's Uncertainty Principle

Werner Heisenberg

It is impossible to determine simultaneously, the exact position and exact momentum of an electron.

$$\Delta X \times \Delta p_x \geq \frac{h}{4\pi}$$

$$\text{or } \Delta X \times m\Delta V_x \geq \frac{h}{4\pi}$$

$$\text{or } \Delta X \times \Delta V_x \geq \frac{h}{4\pi m}$$

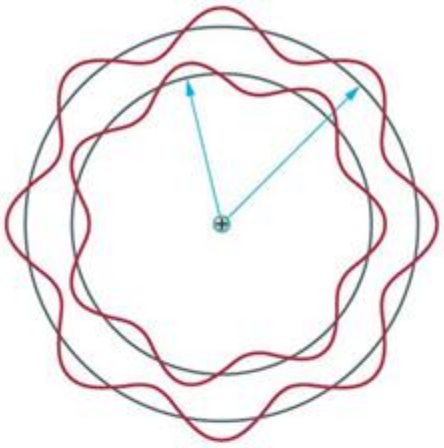
h = Plank's constant

ΔX = Uncertainty in position

Δp_x or ΔV_x = Uncertainty in momentum or velocity

Significance of uncertainty principle

- Rules out Existence of definite paths or trajectories of electrons.



- It is significant only for motion of microscopic objects(e, p, n) and is negligible for that of macroscopic objects (car, bus, ball)

Solved Problems

Q. Uncertainty in position of a hypothetical subatomic particle is 1Å and uncertainty in velocity is $3.3/4\pi \times 10^5$ m/s then calculate the approximate mass of the particle. [$h = 6.6 \times 10^{-34}$ Js]

Sol.

$$\Delta X \times m\Delta V_x \geq \frac{h}{4\pi}$$

$$\Delta V_x = \frac{3.3}{4\pi} \times 10^5 \text{ m/s}$$

$$m \geq \frac{h}{\Delta X \times \Delta V_x \times 4\pi}$$

$$\Delta X = 1 \times 10^{-10} \text{m}$$

$$h = 6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}$$

$$m \geq \frac{6.626 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}}{1 \times 10^{-10}\text{m} \times \frac{3.3}{4\pi} \times 10^5 \text{m/s} \times 4\pi}$$

$$m \geq 2.007 \times 10^{-29} \text{ Kg}$$

Concept Test

Q. An electron is confined to a region of width 5.00×10^{-11} m, which is its uncertainty in position. Estimate the minimum uncertainty in its momentum.

Q. An electron is confined to a region of width 5.00x 10^{-11} m, which is its uncertainty in position. Estimate the minimum uncertainty in its momentum.

Sol.

$$\Delta X \times \Delta p_x \geq \frac{h}{4\pi}$$

$$\Delta X = 5.00 \times 10^{-11} m$$

$$\Delta p_x \geq \frac{h}{\Delta X \times 4\pi}$$

$$h = 6.626 \times 10^{-34} \text{ kg m}^2 s^{-1}$$

$$\Delta p_x \geq \frac{6.626 \times 10^{-34}}{5.00 \times 10^{-11} \times 4\pi}$$

$$\Delta p_x \geq 1.055 \times 10^{-24} \text{ Kg m s}^{-1}$$

Summary

Dual Behaviour of Matter

Louis de Broglie Wavelength $\lambda = \frac{h}{p} = \frac{h}{mv}$

Heisenberg's Uncertainty Principle

It is impossible to determine simultaneously, the exact position and exact momentum of an electron.

$$\Delta X \times \Delta p_x \geq \frac{h}{4\pi}$$

Significance of uncertainty principle

2.5 Towards Quantum Mechanical Model of the Atom

Reference question:

NCERT exercise question: 2.20, 2.21, 2.57, 2.58, 2.59, 2.60, 2.61

2.6

Quantum Mechanical Model of an Atom

2.6 Quantum Mechanical Model of an Atom

Learning objectives

- Quantum Mechanics
- Quantum numbers
- Shape of atomic orbitals
- Effective Nuclear Charge
- Electronic configuration

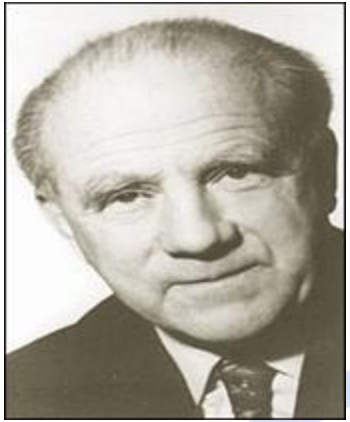
2.6.1 Quantum Mechanics

Classical mechanics v/s Quantum mechanics

Classical mechanics	Quantum mechanics
Dual behavior of matter	Based on wave properties of matter
Behavior of microscopic particles such as protons and electrons	Explain behavior of microscopic particles

Quantum mechanics

It deals with dual behavior of matter both wave like and particle like properties.



Werner Heisenberg

Erwin Schrodinger

Schrodinger equation

$$\boxed{\hat{H}\varphi = E\varphi}$$

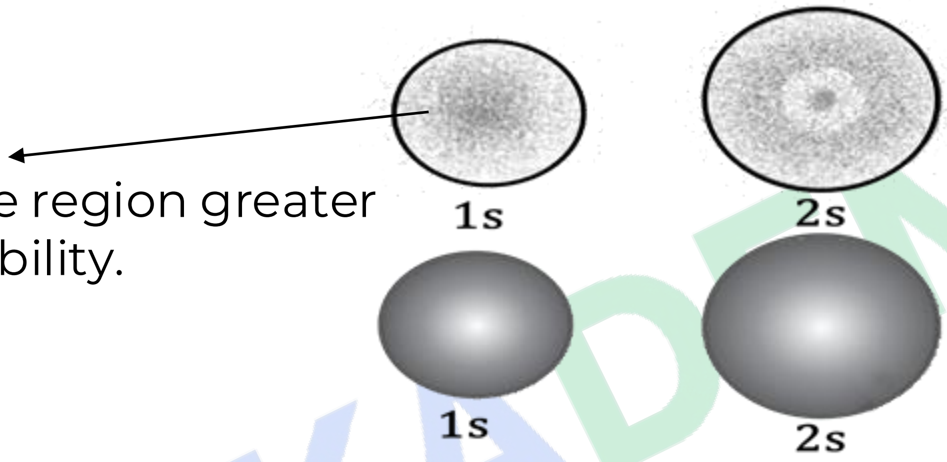
φ → Wave function

$\hat{H}$ → Hamiltonian operator

E → Total energy

$|\varphi|^2$ → **Probability of finding electron density**

Boundary Surface diagram



Darker the region greater the probability.

Drawn for orbital for which $|\varphi|^2$ is constant.

Encloses the region with probability density of more than 90%.

Why do we not draw a boundary surface diagram, which bounds a region in which the probability of finding the electron is 100%?

At any distance from the nucleus, the probability density of finding an electron is never zero.

Orbital→ Region in space where there is a probability of finding an electron is maximum

Node→ Region in space where there is a probability of finding an electron is zero

2.6.2 Quantum Numbers

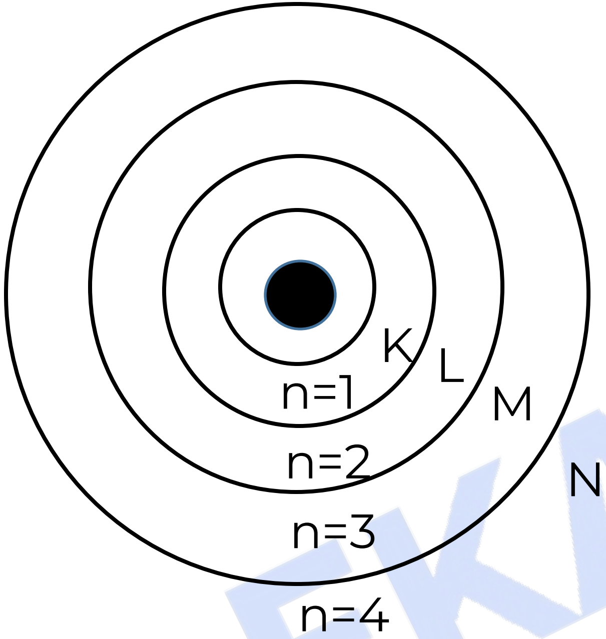
Quantum numbers

It determines size, shape and orientation of orbital

Quantum numbers

- Principle Quantum number
- Azimuthal Quantum number
- Magnetic Quantum number
- Spin Quantum number

Principle Quantum number(n)



$n = 1 \;\; 2 \;\; 3 \;\; 4 \;\; 5$
$\text{Shell} = K \;\; L \;\; M \;\; N \;\; O$

Shell number

Size and energy of orbital

Electrons having same value of n belong to same shell.

$n = 1$ means it belongs to 1^{st} shell.
$n = 2$ means it belongs to 2^{nd} shell.

For a given value of n number of orbital $= n^2$

Azimuthal quantum number(l)

Orbital angular momentum/ subsidiary quantum number

Subshell

$$\text{Value of } l = 0 \quad 1 \quad 2 \quad 3 \quad 4$$
$$\text{Notation for subshell} = s \quad p \quad d \quad f$$

Shape of the orbital

Number of subshell in a given shell is equal to value of n.

Example: For $n = 2$ number of subshell equal to n.

n	l	Subshell notation
1	0	1s
2	0	2s
2	1	2p

$$\text{value of } l \rightarrow 0 \text{ to } (n-1)$$

Subshell

Same n → same shell

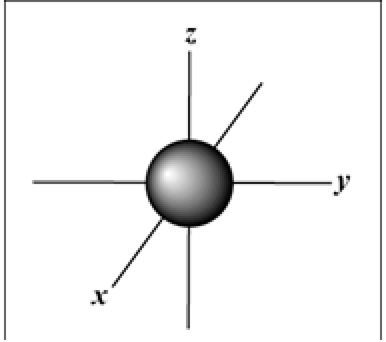
Within the same shell.

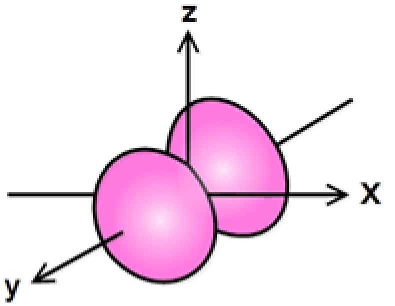
s→spherical

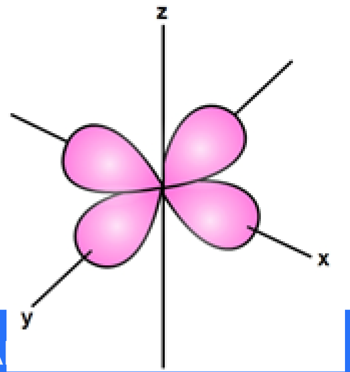
p→collection of dumbbell-shape

d→collection of double dumbbell-shape

Boundary surface plots which are similar are said to be same subshell.

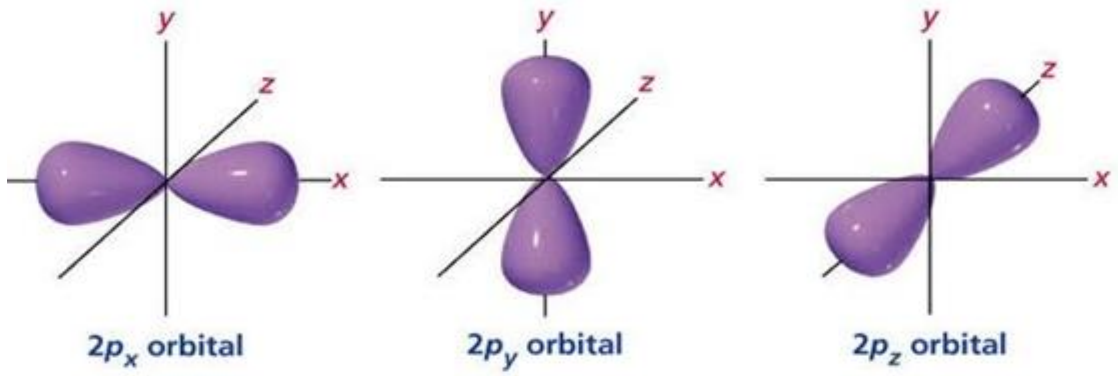






Magnetic Quantum Number (m_l)

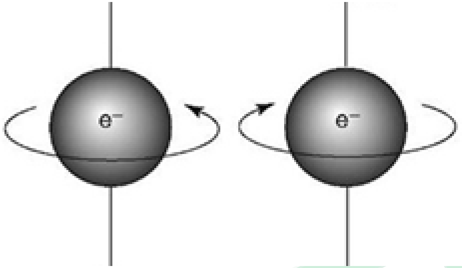
Orientation of the orbital with respect to standard set of co-ordinate axis.



For any value of subshell (l), $2l + 1$ values of m are possible

$$\boxed{\text{Value of } m_l = -l \text{ to } +l}$$

Spin Quantum Number (m_s)



Describes the angular momentum of electron

Vector quantity having both magnitude and direction.

Represented by arrow
↑ –(spin up)
↓ –(spin down)

Spin of electron $+\frac{1}{2}$ and $-\frac{1}{2}$ are said to have opposite spins.

Solved Problem

Q. How many subshells are associated with $n = 4$?

Sol.

For $n = 4$, $l = 0, 1, 2, 3$

So, 4 subshell s, p, d, f.

Concept Test

Q. Which of the following orbitals are possible? $1p$, $2s$, $2p$ and $3f$.

Q. Which of the following orbitals are possible? 1p, 2s, 2p and 3f.

Sol.

For 1p, value of $n = 1, l = 1$, l can be only 0 so 1p not possible

For 2s, value of $n = 2, l = 0$ possible

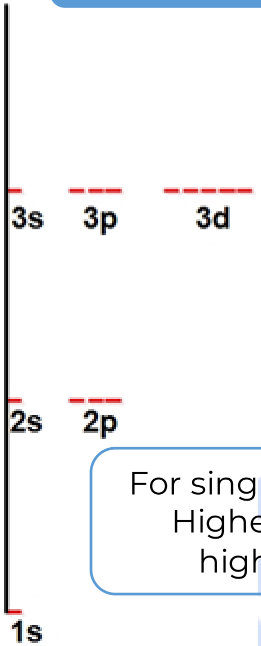
For 2p, value of $n = 2, l = 1$ possible

For 3f, value of $n = 3, l = 3$, l can only be $0, 1, 2$ so not possible

2.6.3
Effective Nuclear Charge

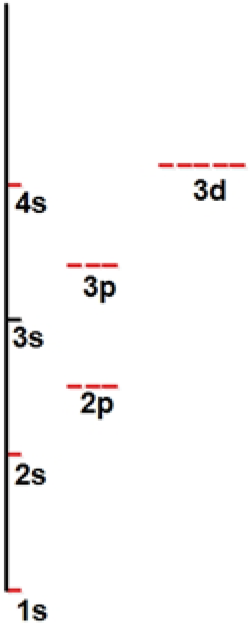
Degenerate orbital

Orbitals having same energy are called degenerate orbitals



For single electron system,
Higher the value of n,
higher the energy

For single electron species energy of orbitals having same quantum number have same energy.



For multi electron system,
Higher the value of (n+l), higher the energy

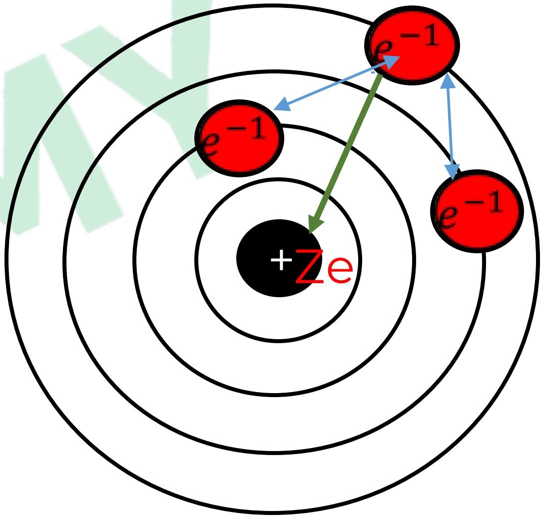
For multi-electron atom orbitals with same quantum number posses different energy

Effective nuclear charge ($Z_{effective}$)

Shielding of outer shell electrons from the nucleus by the inner shell electrons.

The net positive charge experienced by the outer electrons is known as effective nuclear charge.

$$Z_{eff} < Z$$



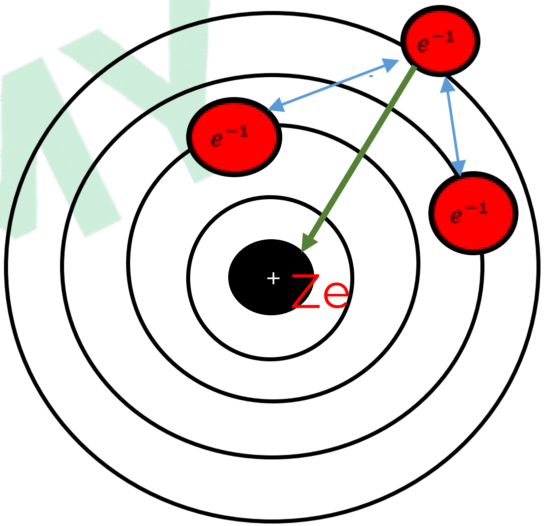
Attraction by nucleus and repulsion by electron in inner shell

Factors affecting Z_{eff}

$n \uparrow \rightarrow$ distance from nucleus $\uparrow \rightarrow$ screening $\uparrow$, $Z_{eff} \uparrow$

$Z_{eff}(1s) > Z_{eff}(2s)$

$Z_{eff}(4p) < Z_{eff}(3p)$



Azimuthal quantum number(l) $\uparrow$ $Z_{eff} \downarrow$

$\Downarrow$

Z_{eff} $s > p > d > f$

$Z_{eff}(4s) > Z_{eff}(4p)$

Atomic number $\uparrow$
Energies of orbitals in the same subshell $\downarrow$

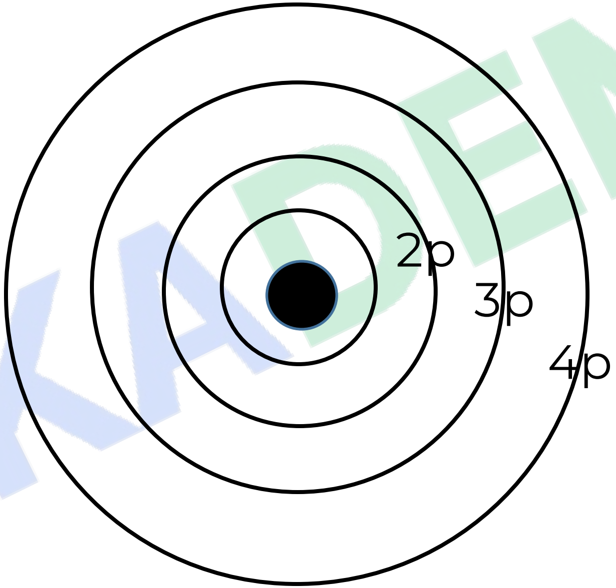
$E_{2s}(H) > E_{2s}(Li) > E_{2s}(Na)$

Concept Test

Q. The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electron in 4p orbital. Which of these electron experiences the lowest effective nuclear charge ?

Q. The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electron in 4p orbital. Which of these electron experiences the lowest effective nuclear charge ?

Sol.

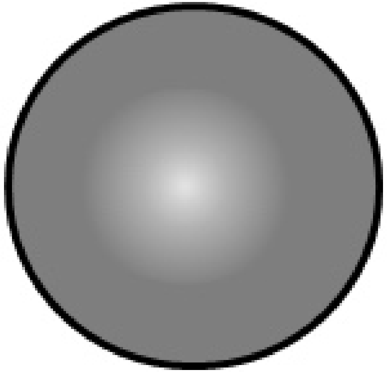


4p will experience lowest effective nuclear charge as it is farthest from the nucleus.

2.6.4
Shape of Atomic Orbital

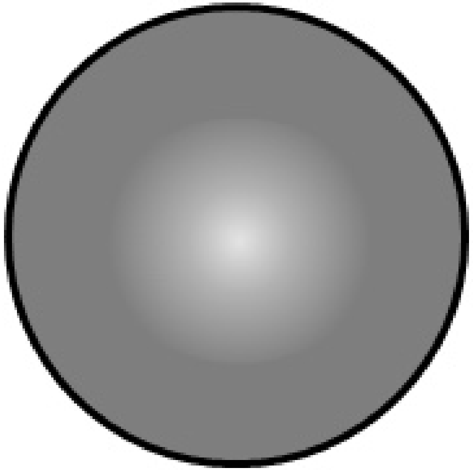
Boundary surface diagram of 1s and 2s orbital.

for $n = 1, l = 0$



1s

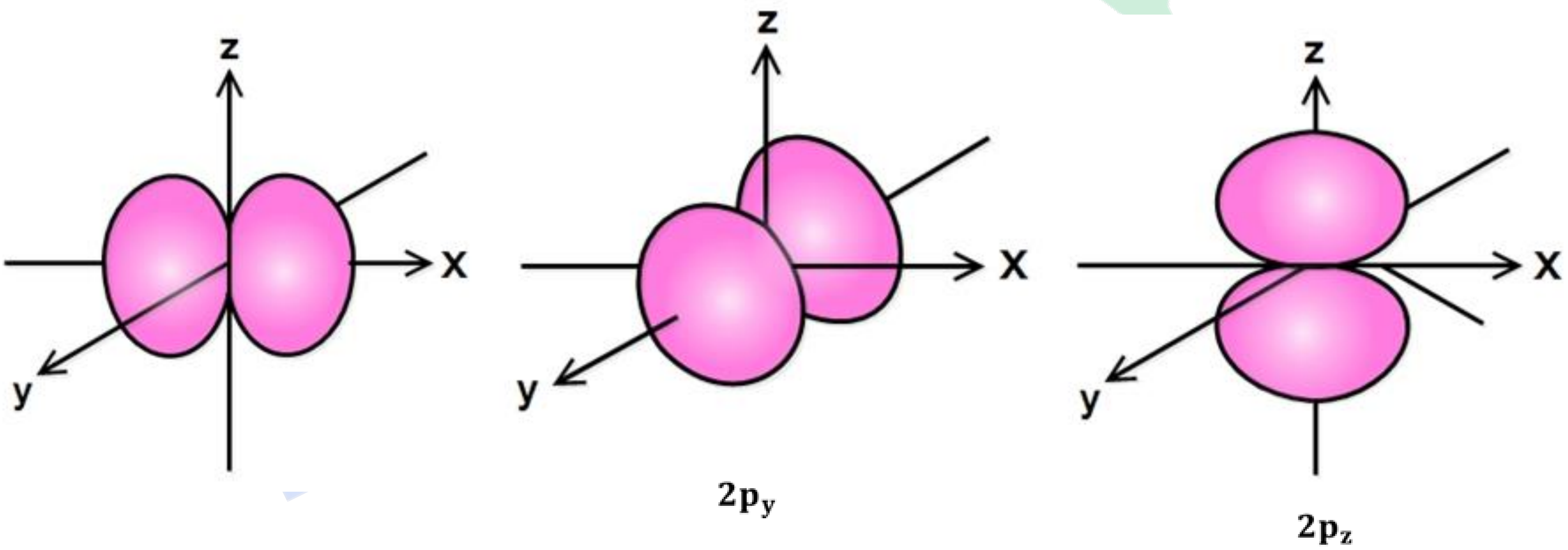
for $n = 2, l = 0$



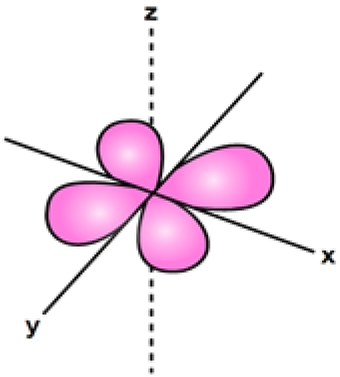
2s

Boundary surface diagram of the three 2p orbital

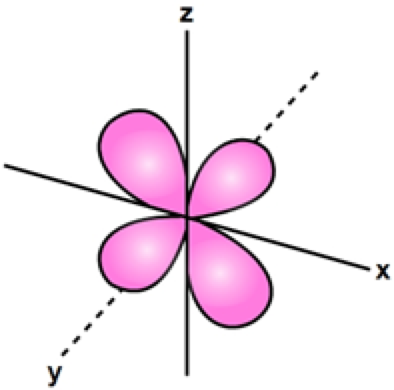
$$\text{for } n = 2, l = 1 \;\; m_l = -1, 0, 1$$



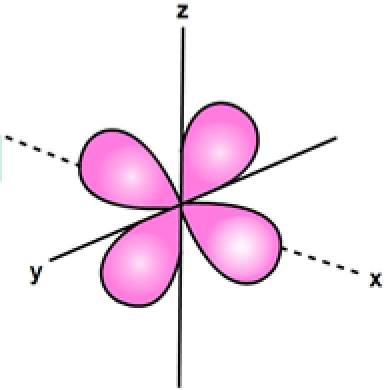
Boundary surface diagram of the five 3d orbital.



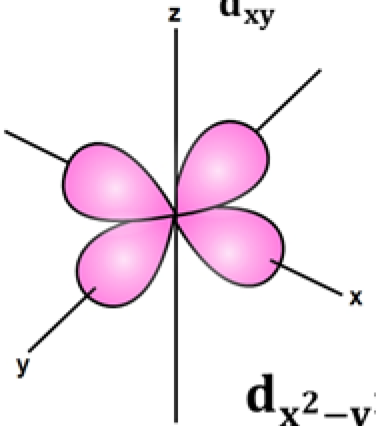
d_{xy}



d_{xz}

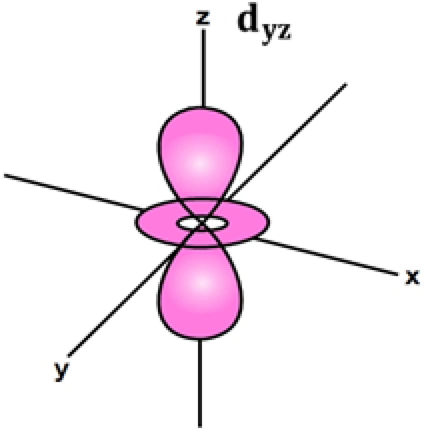


d_{yz}

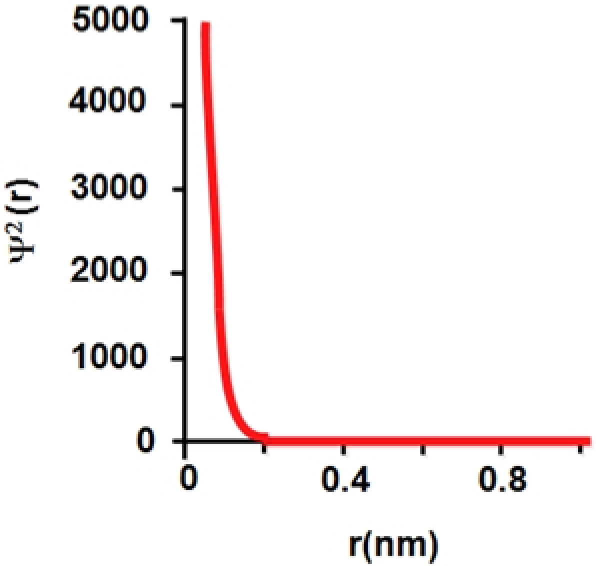


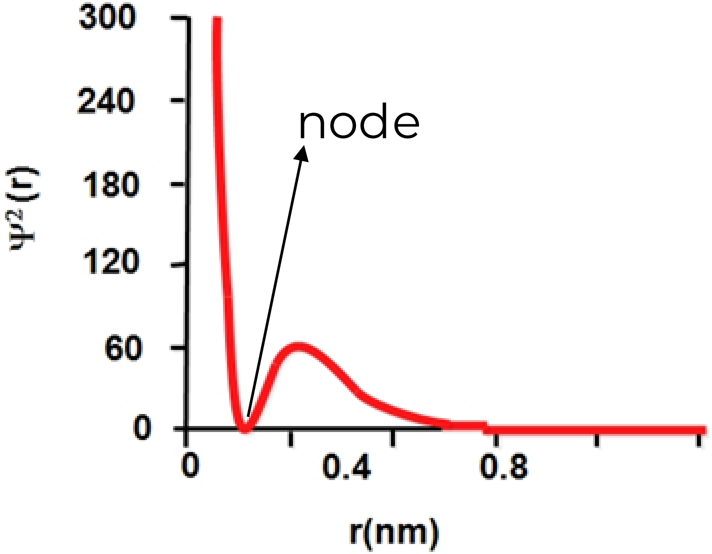
$d_{x^2-y^2}$

For n = 3, l = 2, m_l = -2, -1, 0, +1, +2



Variation of probability density $|\varphi|^2$ as a function of distance, r, of the electron from the nucleus for 1s and 2s orbitals





$$Radial\ node = (n - l - 1)$$
$$Angular\ node/nodal\ plane = 1$$
$$Total = (n - l)$$

2.6.5

Electronic Configuration

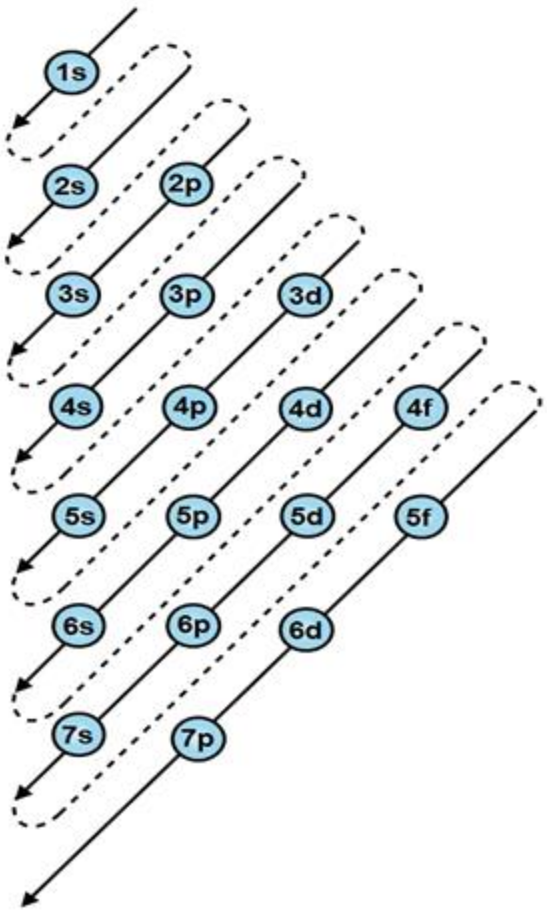
Aufbau principle

In ground states of the atoms, the orbitals are filled in order of their increasing energies value of (n+l).

For same value of (n + l) n ↑ energy ↑

Orbital	Value of n	Value of l	Value of (n+l)	
1s	1	0	1	
2s	2	0	2	
2p	2	1	3	2p(n=2)has lower energy than
3s	3	0	3	3s(n=3)
3p	3	1	4	3p(n=3)has lower energy than
4s	4	0	4	4s(n=4)
3d	3	2	5	3d(n=3)has lower energy than
4p	4	1	5	4p(n-4)

Aufbau principle



Pauli Exclusion Principle

No two electron in an atom can have the same set of four quantum number.

Wolfgang Pauli

		n	l	m	s
2s	Electron 1	2	0	0	
	Electron 2	2	0	0	

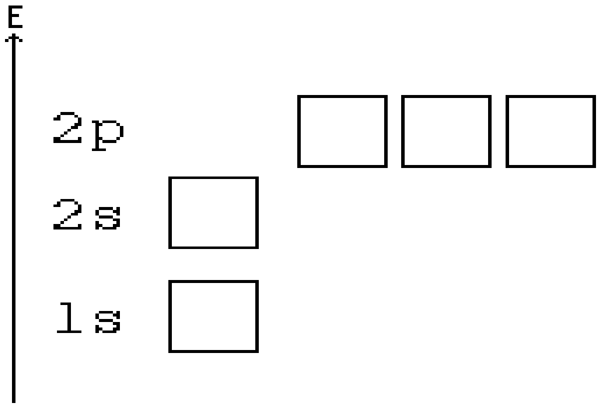
It shows that only two electron may exist in the same orbital and these electrons must have opposite spin.

So maximum number of electrons in a shell with quantum number n is $2n^2$

Hund’ s rule of maximum multiplicity

Pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not take place until each orbital belonging to that subshell has got one electron each i.e., it is singly occupied.

Friedrich Hund



Electronic Configuration

Electronic configuration can be represented by

$s^a p^b d^c$ notation

subshell→respective letter symbol
Quantum number before respective subshell
Electron in subshell→as superscript

Orbital diagram

Orbital of subshell→by box
electron→ arrow (↑)
clockwise spin
arrow (↓) anti-clockwise spin

Example: $Li \rightarrow 1s^2 2s^1$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$

$F \rightarrow 1s^2 2s^2 2p^5$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow}$

Stability of half filled and completely filled subshell

Why actual electronic configuration was different from expected?

Electronic configuration of chromium(24)

$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2 \text{(expected)}$$

$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1 \text{(actual — more stable)}$$

Electronic configuration of copper(29)

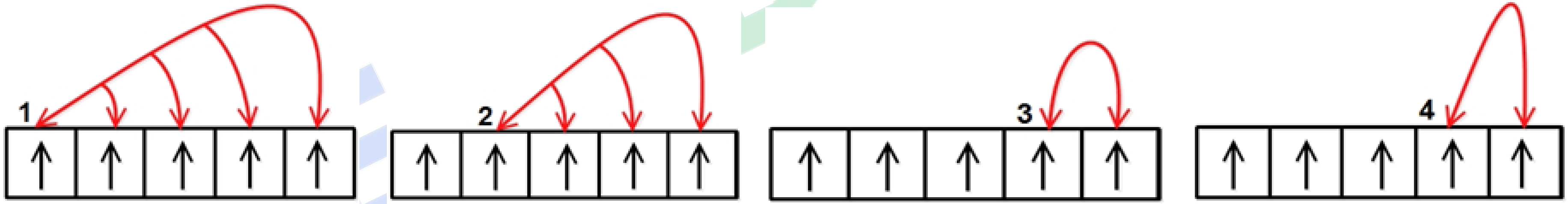
$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2 \text{(expected)}$$

$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1 \text{(actual — more stable)}$$

Causes of Stability of half filled and completely filled subshell

1. Symmetrical distribution of electrons

2. Exchange energy



4 exchange by electron 1

3 exchange by electron 2

2 exchange by electron 3

1 exchange by electron 4

Summary

Schrodinger equation $\hat{H}\varphi = E\varphi$

Quantum numbers $(\mathrm{n}, \mathrm{l}, \mathrm{m_l}, \mathrm{m_s})$
l varies from 0 to n-1
$\mathrm{m_l}$ varies from –l to +l
$\mathrm{m_s} = -\frac{1}{2} \text{ or } +\frac{1}{2}.$
Shape of s, p and d orbitals.

Aufbau principle: (n+l) value↑ energy↑

Pauli exclusion principle: only two electrons can exist in same orbital having opposite spin.

Hund's rule: until each orbital of same subshell is not singly occupied pairing of electron does not takes place

Half filled and full filled orbitals are more stable.

2.6 Quantum Mechanics

Reference question:

NCERT exercise question: 2.28, 2.29, 2.30, 2.31, 2.62, 2.63, 2.64, 2.65, 2.66, 2.67

Thank You