Std-XI science

Unit 2:

STRUCTURE OF ATOM

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Discovery of Electron

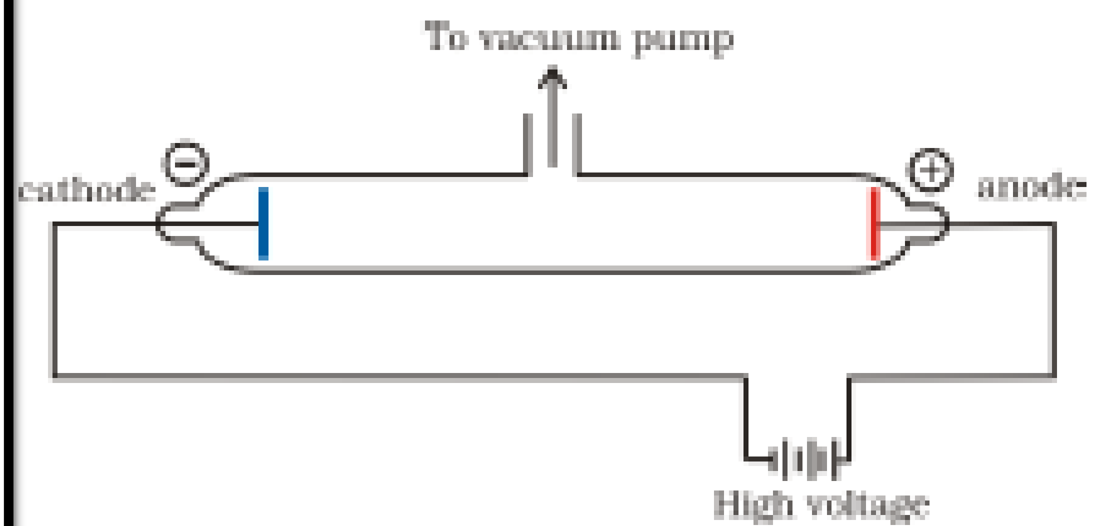


Fig. 2.1(a) A cathode ray discharge tube

Discovery of Electron

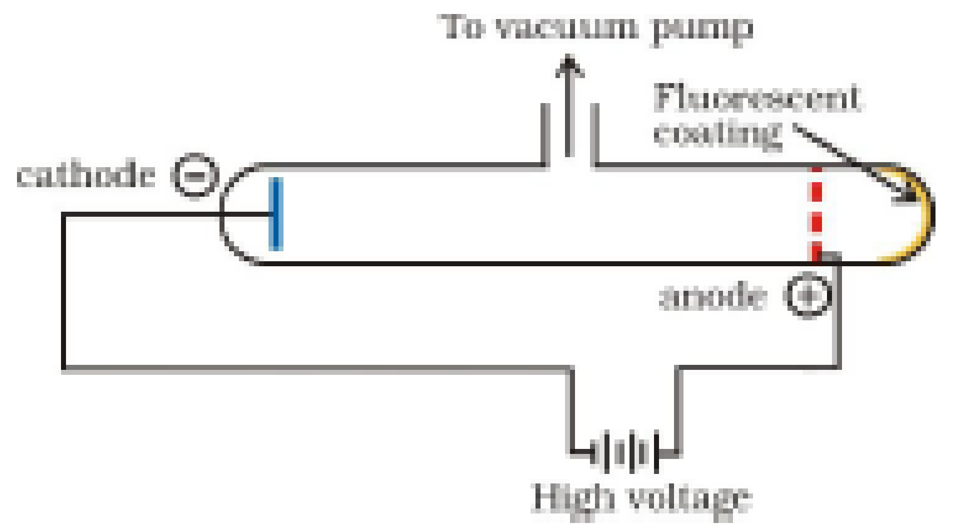


Fig. 2.1(b) A cathode ray discharge tube with perforated anode

Properties of Cathode Rays

(i) The cathode rays start from cathode and move towards the anode.
(ii) 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
(iii) In the absence of electrical or magnetic field, these rays travel in straight lines

Properties of Cathode Rays……contd

(iv) 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.

(v) The characteristics of cathode rays (electrons) do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube.

Thus, we can conclude that electrons are basic constituent of all the atoms

The results of these experiments are summarized below.

- The cathode rays start from cathode and move towards the anode.

- (ii) 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..

Cathode ray discharge Tube

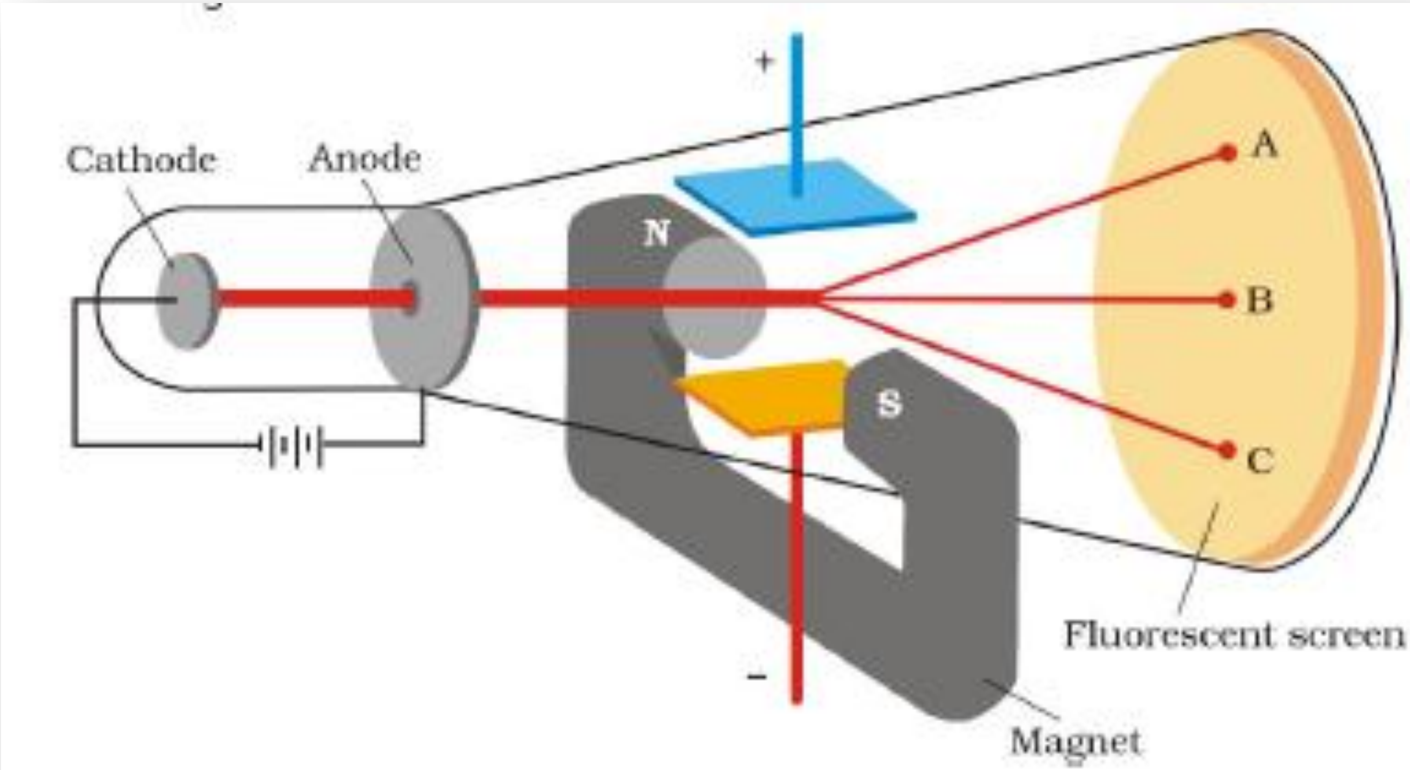


Fig. 2.2 The apparatus to determine the charge to the mass ratio of electron

The results of these experiments are summarized below………contd

- 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
- In the absence of electrical or magnetic field, these rays travel in straight lines

Cathode ray discharge Tube

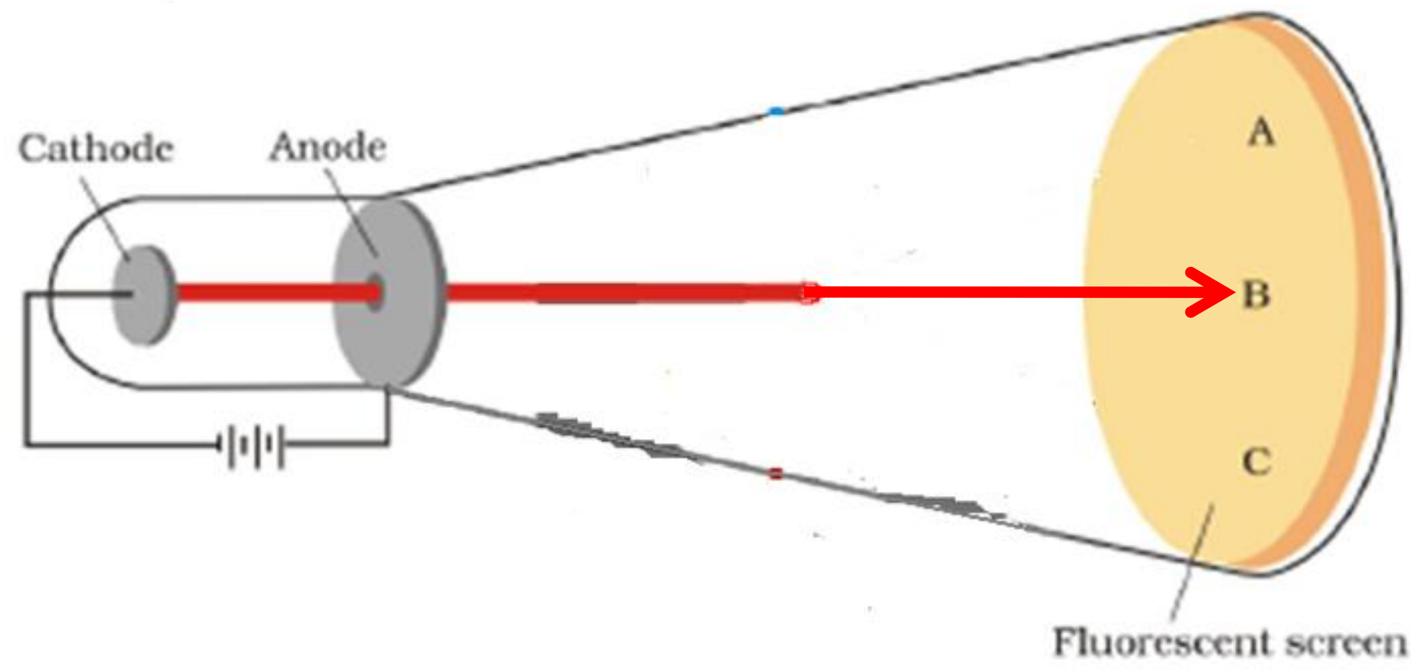


Fig. 2.2 The apparatus to determine the charge to the mass ratio of electron

The results of these experiments are summarized below………contd

- 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.

Effect of Electric field

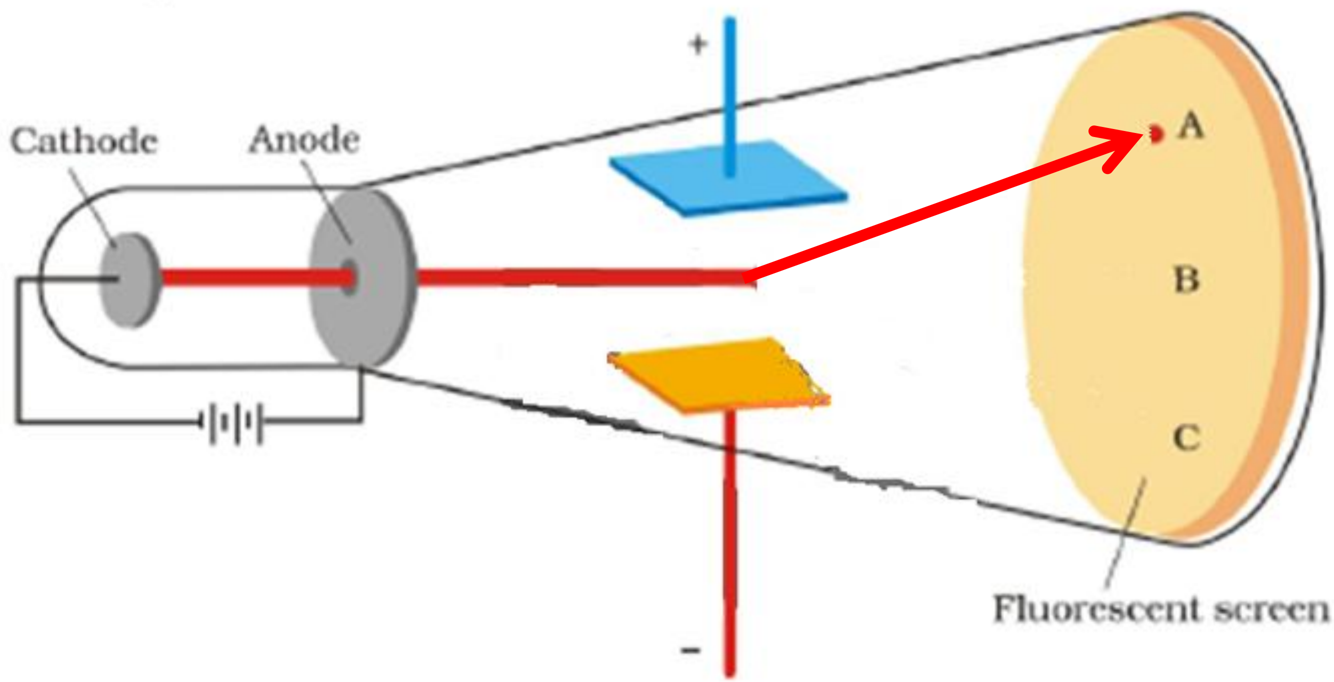


Fig. 2.2 The apparatus to determine the charge to the mass ratio of electron

Effect of Magnetic field

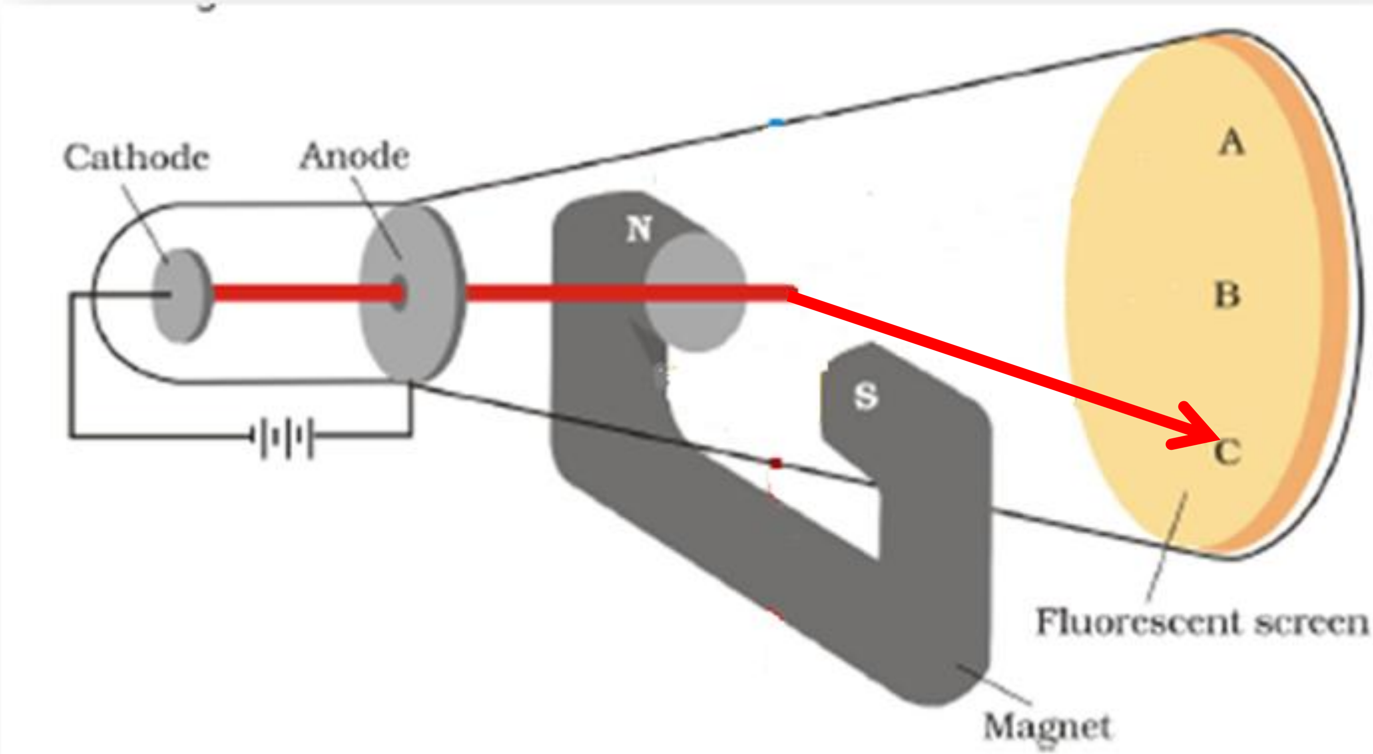


Fig. 2.2 The apparatus to determine the charge to the mass ratio of electron

The results of these experiments are summarized below………contd

- The characteristics of cathode rays (electrons) do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube.

- Thus, *we can conclude that electrons are basic constituent of all the atoms*.

Charge to Mass Ratio of Electron (e/m Ratio)

- In 1897, British physicist J.J. Thomson measured the ratio of electrical charge (e) to the mass of electron (m_e) by using cathode ray tube and applying electrical and magnetic field perpendicular to each other as well as to the path of electrons

Cathode ray discharge Tube

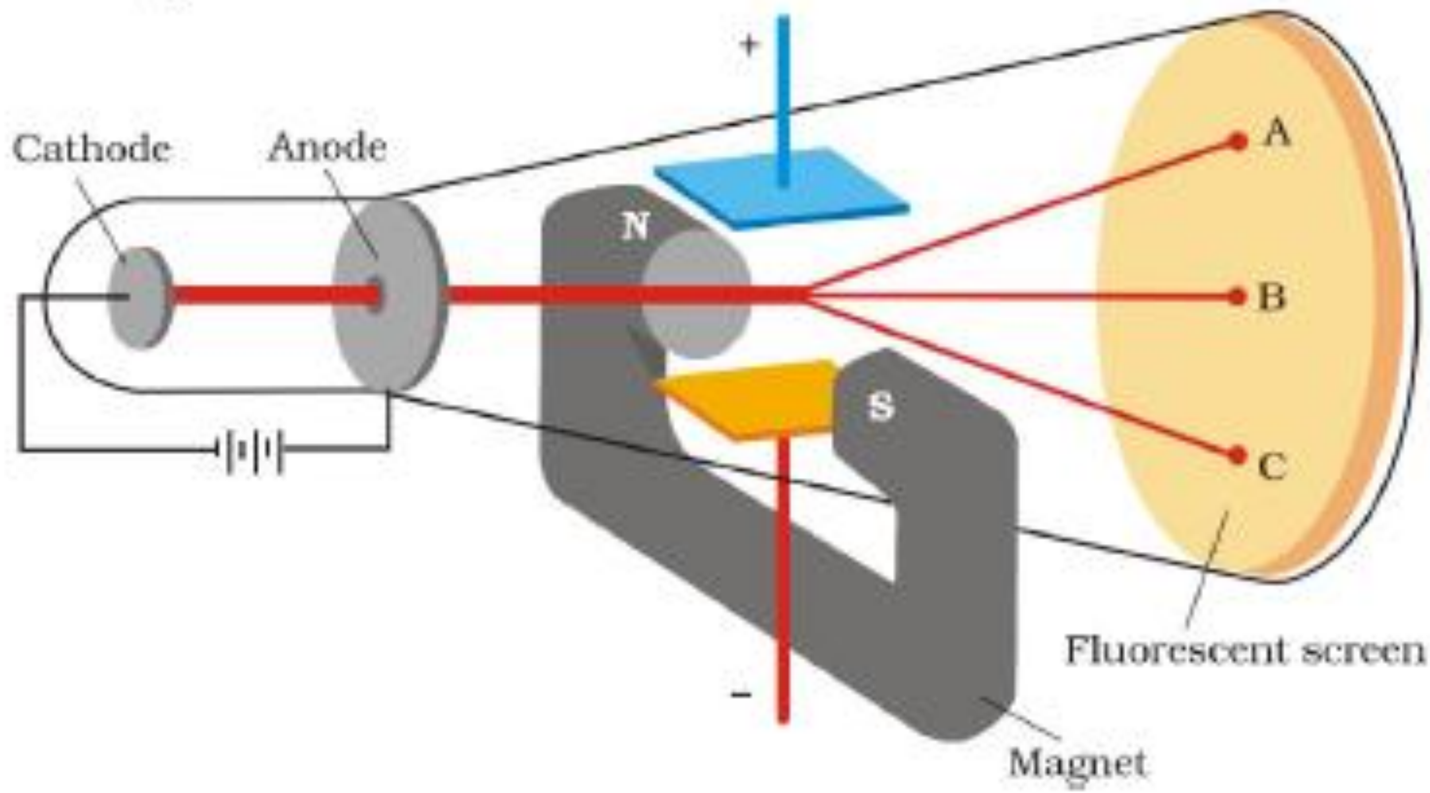


Fig. 2.2 The apparatus to determine the charge to the mass ratio of electron

- Thomson argued that the amount of deviation of the particles from their path in the presence of electrical or magnetic field depends upon:

1. the magnitude of the negative charge on the particle, greater the magnitude of the charge on the particle, greater is the interaction with the electric or magnetic field and thus greater is the deflection.

2. the mass of the particle — lighter the particle, greater the deflection.

3. the strength of the electrical or magnetic field — the deflection of electrons from its original path increases with the increase in the voltage across the electrodes, or the strength of the magnetic field.

- When only electric field is applied, the electrons deviate from their path and hit the cathode ray tube at point A.

- Similarly when only magnetic field is applied, electron strikes the cathode ray tube at point C.

- By carefully balancing the electrical and magnetic field strength, it is possible to bring back the electron to the path followed as in the absence of electric or magnetic field and they hit the screen at point B.

Charge to Mass Ratio of Electron
(e/m Ratio)

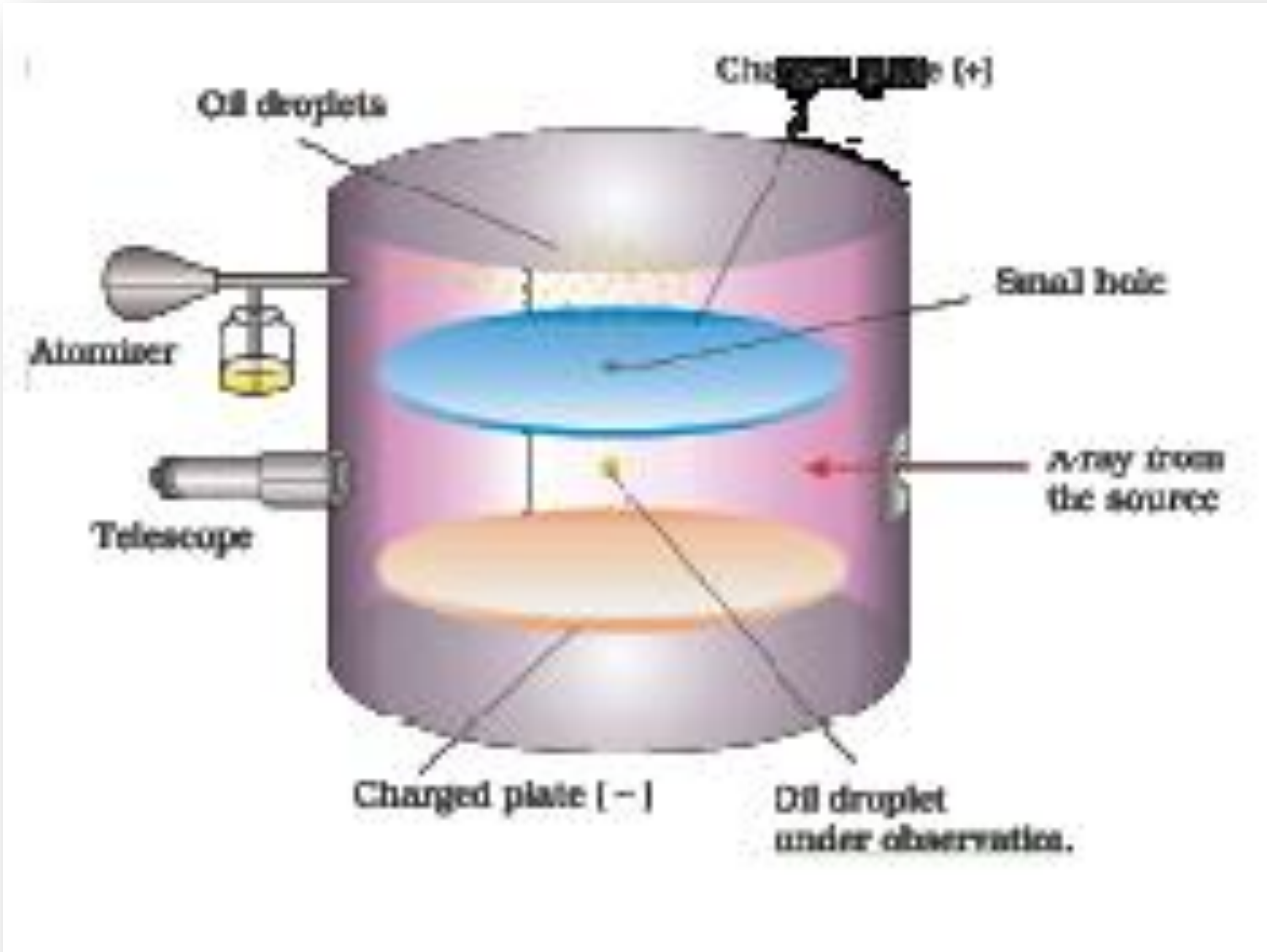
$$\frac{e}{m_e} = 1.758820 \times 10^{11}\ C\ kg^{-1}$$

Where m_e is the mass of the electron in kg and e is the magnitude of the charge on the electron in coulomb (C).

Charge on the Electron

- R.A. Millikan (1868-1953) devised a method known as oil drop experiment (1906-14), to determine the charge on the electrons.
- He found that the charge on the electron to be -1.6×10^{-19} C.
- The present accepted value of electrical charge is
- -1.6022×10^{-19} C.

Millikan's Oil Drop Method



Mass of the Electron

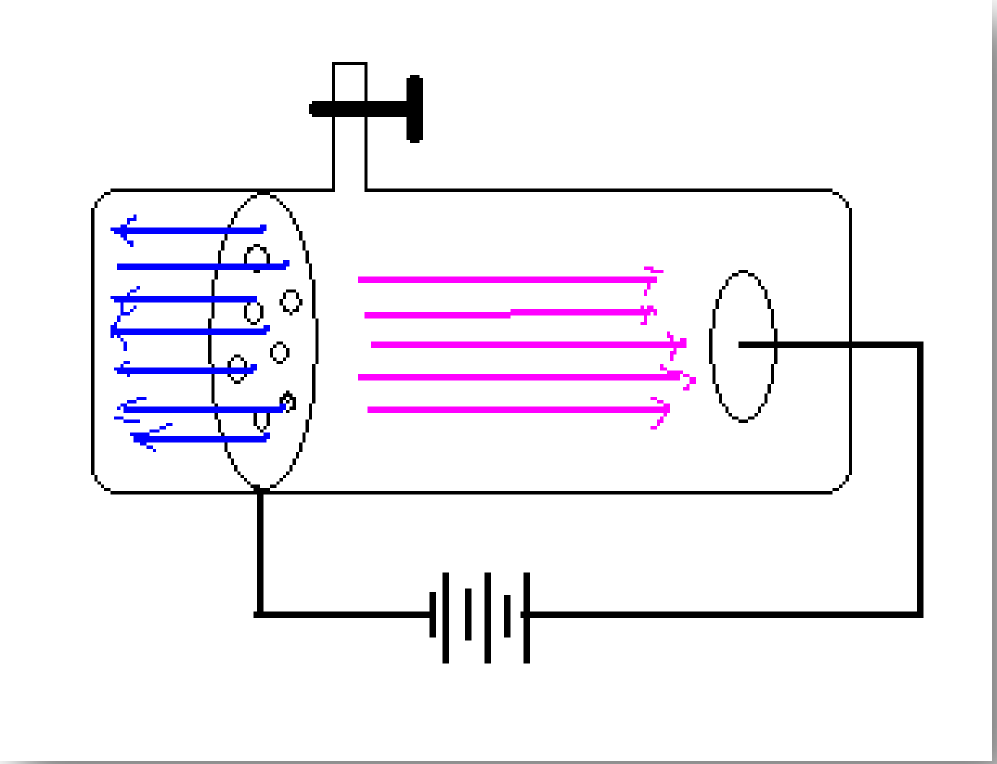
- The mass of the electron (m_e) was determined by combining these results with Thomson's value of e/me ratio.

$$m_e = \frac{e}{e/m_e} = \frac{1.6022 \times 10^{-19} C}{1.758820 \times 10^{11} C\ kg^{-1}}$$

- The mass of the electron (m_e)

- $= 9.1094 \times 10^{-31}$ kg

Discovery of Proton & Neutron



Properties of Canal rays(anode rays)

1. They are deflected by electric field towards negatively charged electric plate indicating that they are positively charged.

2. They get deflected by Magnetic field

Properties of Canal rays(anode rays)

- unlike cathode rays, the positively charged particles depend 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 is found to depend on the gas from which these originate.

Proton

- The smallest and lightest positive ion was obtained from hydrogen and was called proton.
- This positively charged particle was characterised in 1919.

Neutrons

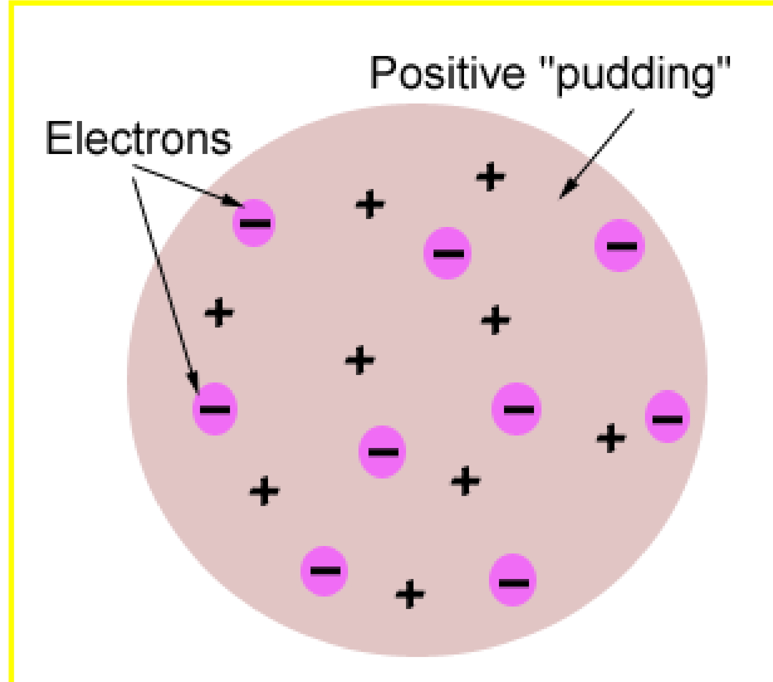
- Later, a need was felt for the presence of electrically neutral particle as one of the constituent of atom.
- These particles were discovered by Chadwick (1932) by bombarding a thin sheet of beryllium by α-particles.
- When electrically neutral particles having a mass slightly greater than that of the protons was emitted.
- He named these particles as neutrons.

Subatomic Particles

Particle	Charge	Mass (g)	Location
Electron (e^-)	-1	9.11×10^{-28}	Electron cloud
Proton (p^+)	+1	1.67×10^{-24}	Nucleus
Neutron (n^o)	0	1.67×10^{-24}	Nucleus

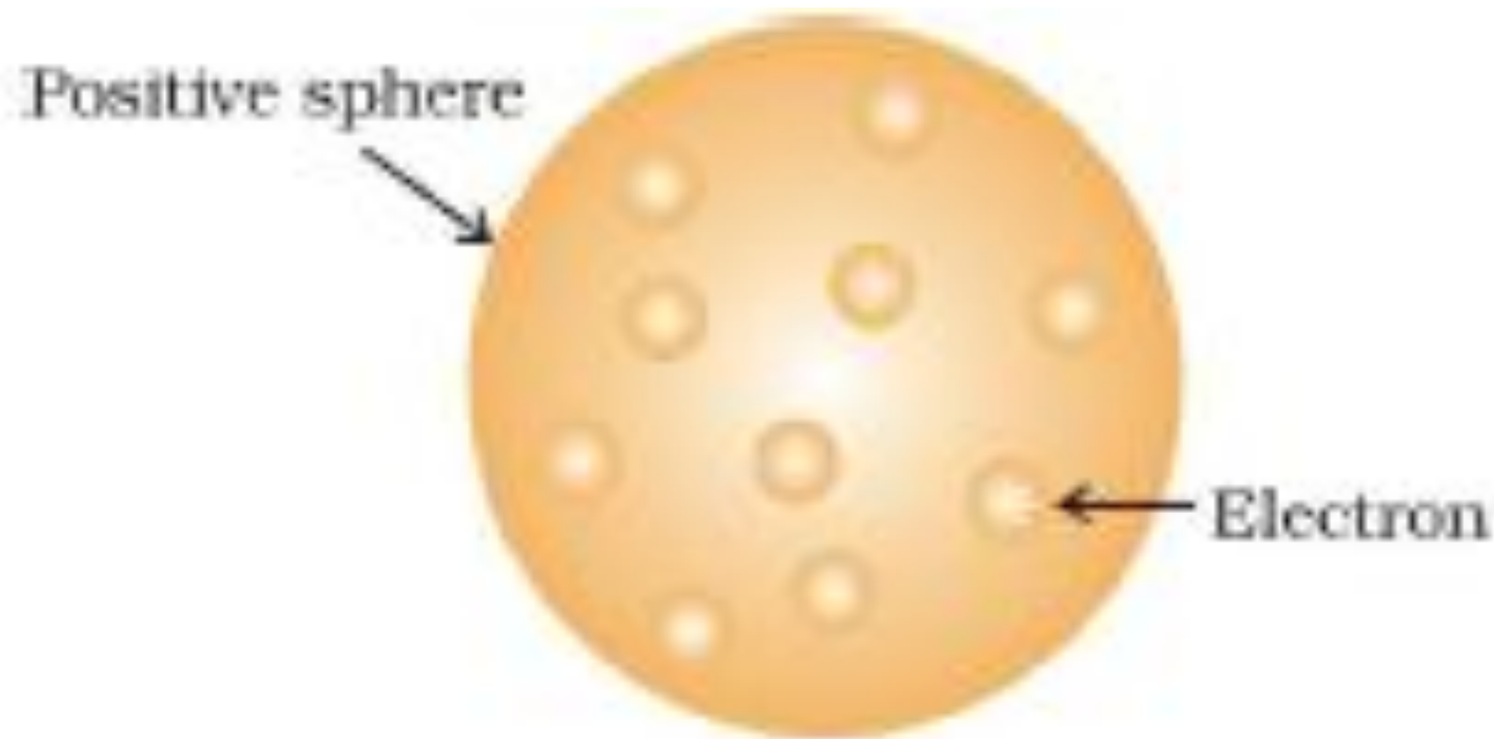
Thomson's Atomic Model

J. J. Thomson



Thomson believed that the electrons were like plums embedded in a positively charged "pudding," thus it was called the "plum pudding" model.

Thomson Model of Atom



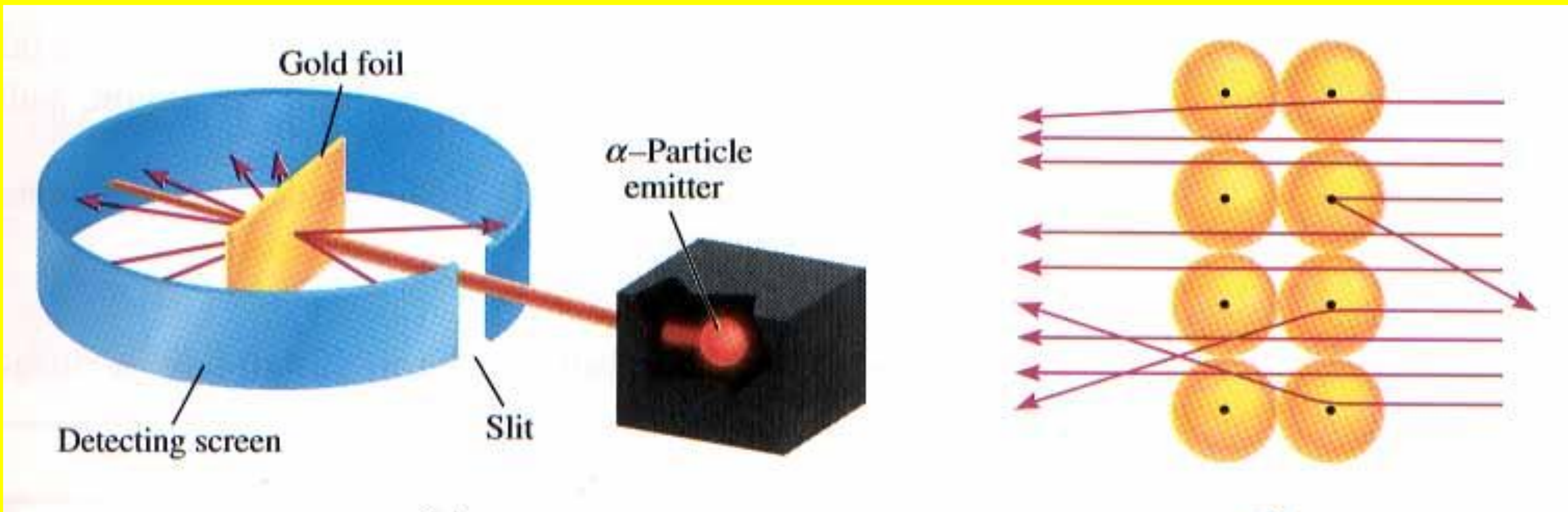
Thomson Model of Atom

- J. J. Thomson, in 1898, proposed that an atom possesses a spherical shape (radius approximately 10^{-10} m) in which the positive charge is uniformly distributed.
- The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement
- plum pudding, raisin pudding or watermelon.

Features of Thomson Model of Atom

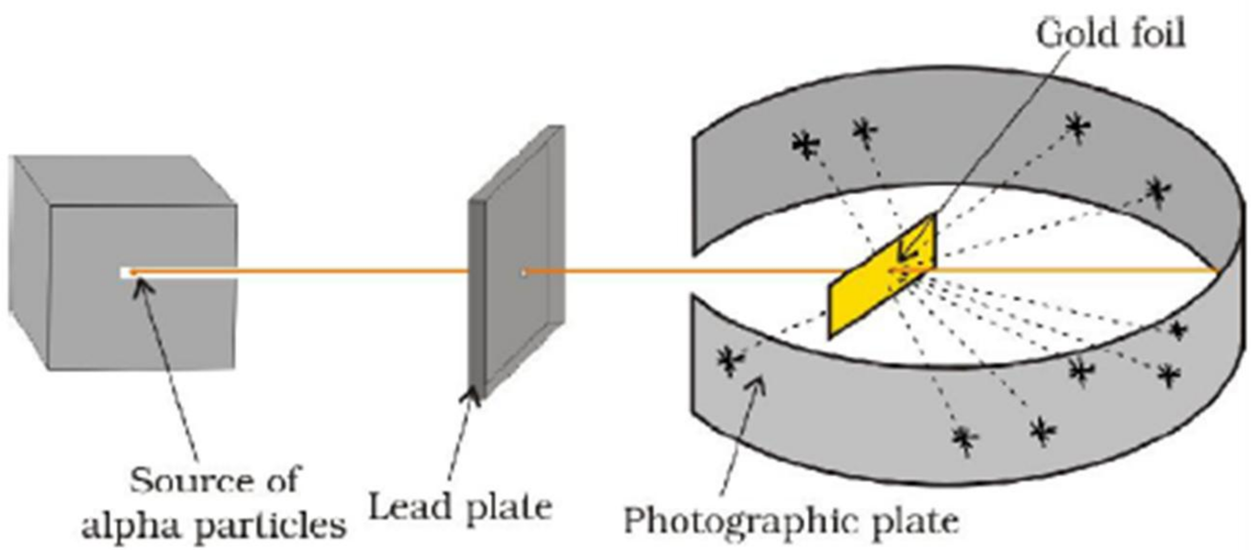
- An important feature of this model is that the mass of the atom is assumed to be uniformly distributed over the atom.
- Although this model was able to explain the overall neutrality of the atom,
- but was not consistent with the results of later experiments.

Ernest Rutherford's Gold Foil Experiment - 1911



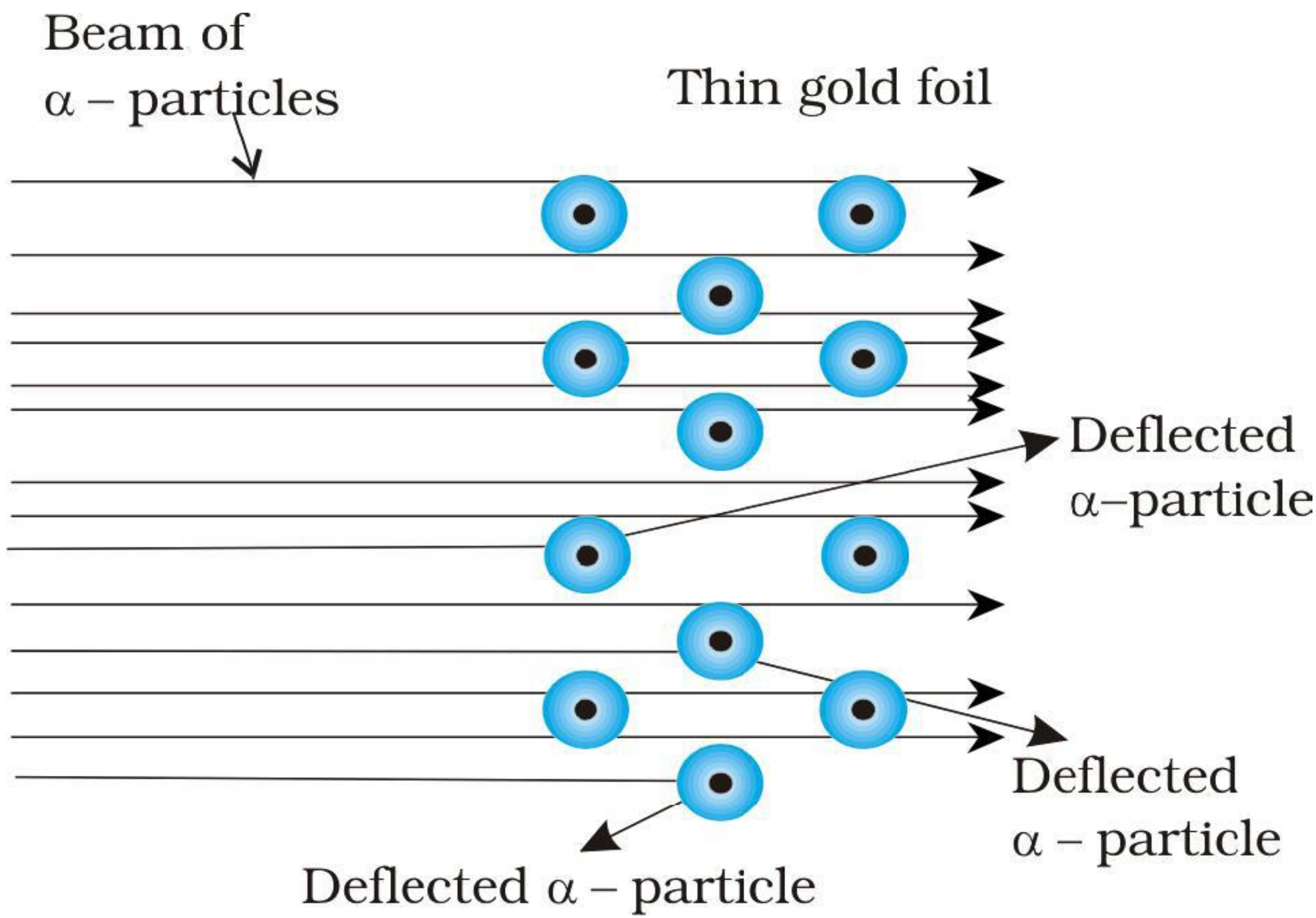
- Alpha particles are helium nuclei - The alpha particles were fired at a thin sheet of gold foil
- Particles that hit on the detecting screen (film) are recorded

Rutherford's Nuclear Model of Atom



A. Rutherford's scattering experiment

Schematic molecular view of the gold foil



Observations

- (i) most of the $\alpha-$ particles passed through the gold foil undeflected.
- (ii) a small fraction of the $\alpha-$particles was deflected by small angles.
- (iii) a very few $\alpha-$ particles (~1 in 20,000) bounced back, that is, were deflected by nearly 180°.

Conclusions

- (i) Most of the space in the atom is empty as most of the α–particles passed through the foil undeflected.
- (ii) A few positively charged α– particles were deflected. The deflection must be due to enormous repulsive force showing that the positive charge of the atom is not spread throughout the atom as Thomson had presumed.
- The positive charge has to be concentrated in a very small volume that repelled and deflected the positively charged α– particles.

- (iii) Calculations by Rutherford showed that the volume occupied by the nucleus is negligibly small as compared to the total volume of the atom.

The Rutherford Atomic Model

- Based on his experimental evidence:
  - The atom is mostly empty space
  - All the positive charge, and almost all the mass is concentrated in a small area in the center. He called this a "nucleus"
  - The nucleus is composed of protons and neutrons (they *make* the nucleus!)

- The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called orbits.
- Thus, Rutherford's model of atom resembles the solar system in which the nucleus plays the role of sun and the electrons that of revolving planets.
- His model was called a "nuclear model"

Atomic Number and Mass Number

- The number of protons present in the nucleus is equal to atomic number (Z) .

- Atomic number (Z)
  - = number of protons in the nucleus of an atom
  - = number of electrons in a neutral atom

Nucleons

- While the positive charge of the nucleus is due to protons, the mass of the nucleus, due to protons and neutrons.

- protons and neutrons present in the nucleus are collectively known as nucleons.

Mass number (A)

- The total number of nucleons is termed as mass number (A) of the atom.

- Mass number (A)
  - = number of protons (Z)
  - + number of neutrons (n)

Atomic Number

Atomic number (Z) of an element is the number of protons in the nucleus of each atom of that element.

Element	# of protons	Atomic # (Z)
Carbon	6	6
Phosphorus	15	15
Gold	79	79

Mass Number

Mass number is the number of protons and neutrons in the nucleus of an isotope: Mass # = p^+ + n^0

Nuclide	p^+	n^0	e^-	Mass #
Oxygen - 18	8	10	8	18
Arsenic - 75	33	42	33	75
Phosphorus - 31	15	16	15	31

Complete Symbols

- Contain the symbol of the element, the mass number and the atomic number.

Superscript → Mass number

$$\Large X$$

Subscript → Atomic number

Symbols

- Find each of these:

a) number of protons

b) number of neutrons

c) number of electrons

d) Atomic number

e) Mass Number

$^{80}_{35}Br$

Symbols

- If an element has an atomic number of 34 and a mass number of 78, what is the:

  a) number of protons

  b) number of neutrons

  c) number of electrons

  d) complete symbol

Symbols

- If an element has 91 protons and 140 neutrons what is the

a) Atomic number

b) Mass number

c) number of electrons

d) complete symbol

Symbols

- If an element has 78 electrons and 117 neutrons what is the

a) Atomic number

b) Mass number

c) number of protons

d) complete symbol

Isotopes

- Dalton was wrong about all elements of the same type being identical
- Atoms of the same element *can* have different numbers of neutrons.
- Thus, different mass numbers.
- These are called **isotopes**.

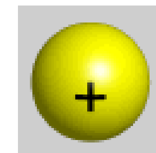
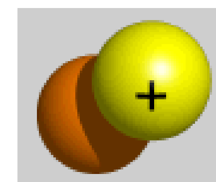
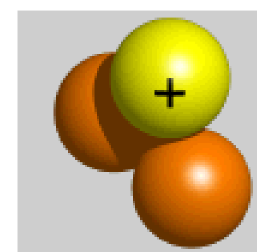
Isotopes

- Frederick Soddy (1877-1956) proposed the idea of isotopes in 1912
- **Isotopes are atoms of the same element having *different masses*, due to varying numbers of neutrons.**
- *Soddy won the Nobel Prize in Chemistry in 1921 for his work with isotopes and radioactive materials*.

Naming Isotopes

- We can also put the mass number *after* the name of the element:
  - carbon-12
  - carbon-14
  - uranium-235

Isotopes are atoms of the same element having *different masses*, due to varying numbers of neutrons.

Isotope	Protons	Electrons	Neutrons	Nucleus
Hydrogen–1 (protium)	1	1	0	
Hydrogen-2 (deuterium)	1	1	1	
Hydrogen-3 (tritium)	1	1	2	

Isobars

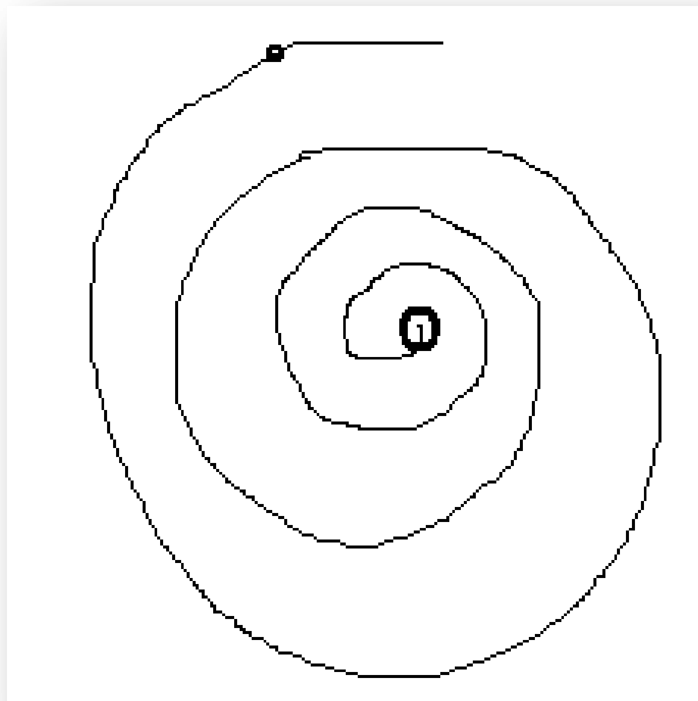
- Isobars are the atoms with same mass number but different atomic number
- example, carbon with Atomic No6 C and mass No14 & Nitrogen with Atomic No7 and mass No14 .

Isoelectronic

i.e., those having the same number of electrons

Drawbacks of Rutherford Model

- Rutherford model tailed in view of electromagnetic theory given by Maxwell.



Drawbacks of Rutherford Model

- *Rutherford's model failed to explain stability of atoms.*
- *Rutherford's model also failed to explain the existence of certain definite lines in the hydrogen spectrum.*

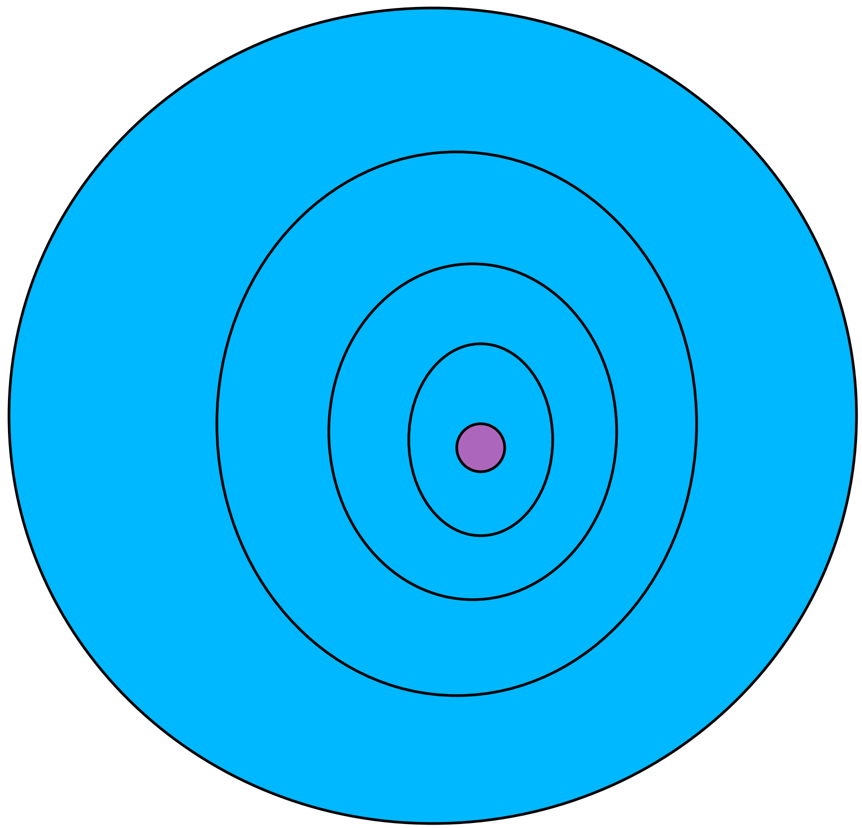
BOHR'S MODEL FOR HYDROGEN ATOM

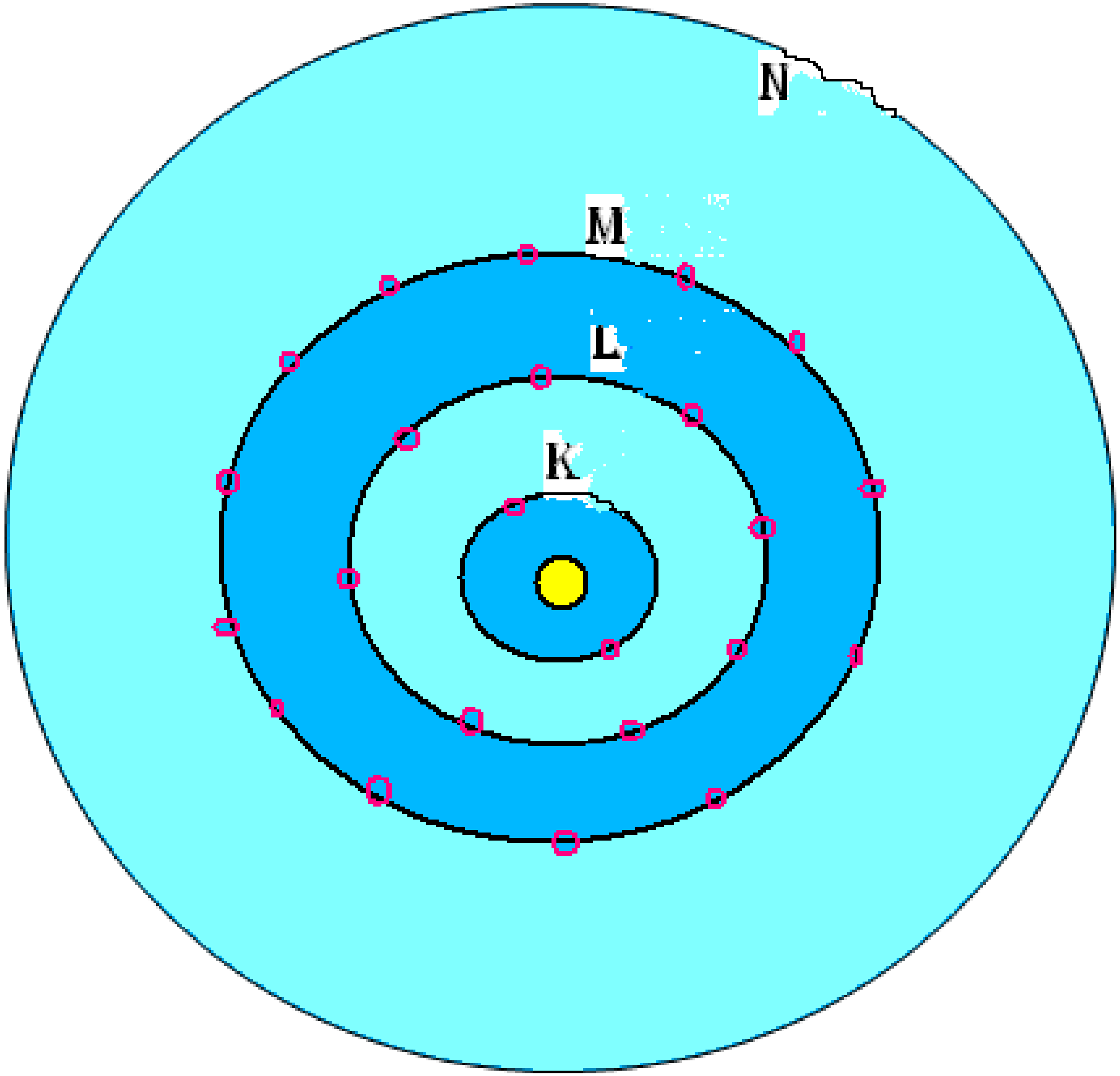
- Neils Bohr (1913) was the first to explain quantitatively the general features of hydrogen atom structure and its spectrum.
- Bohr's model for hydrogen atom is based on the following postulates:

Postulates Bohr's model

- The electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy.
- These paths are called orbits, stationary states or allowed energy states.
- These orbits are arranged concentrically around the nucleus

Bohr's model of an Atom

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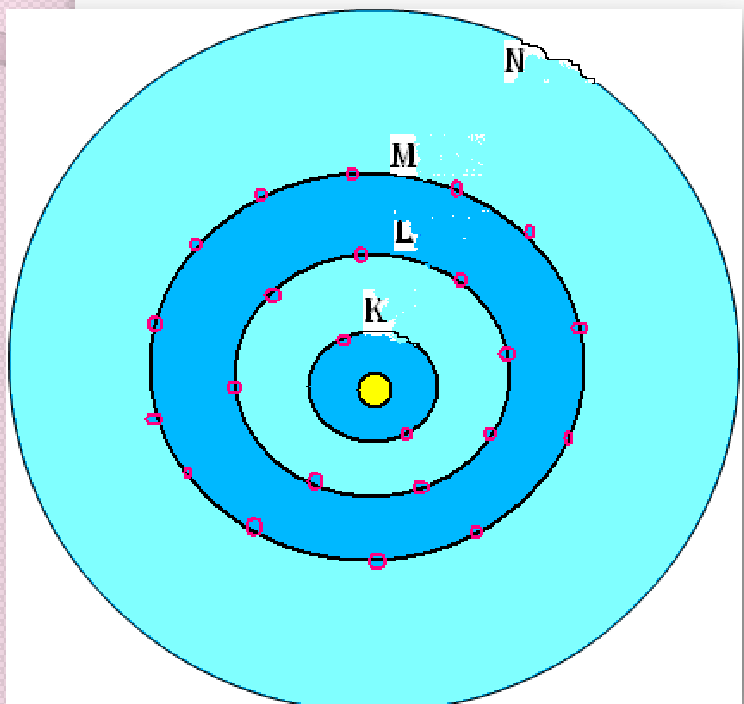
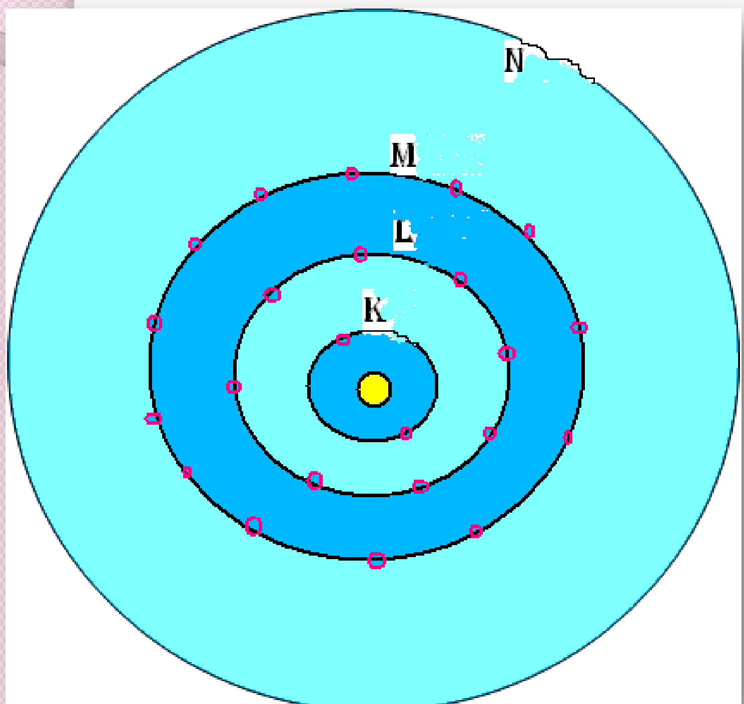
Postulates Bohr's model



- The energy of an electron in the orbit does not change with time. However, the electron will move from a lower stationary state to a higher stationary state when required amount of energy is absorbed by the electron or energy is emitted when electron moves from higher stationary state to lower stationary state

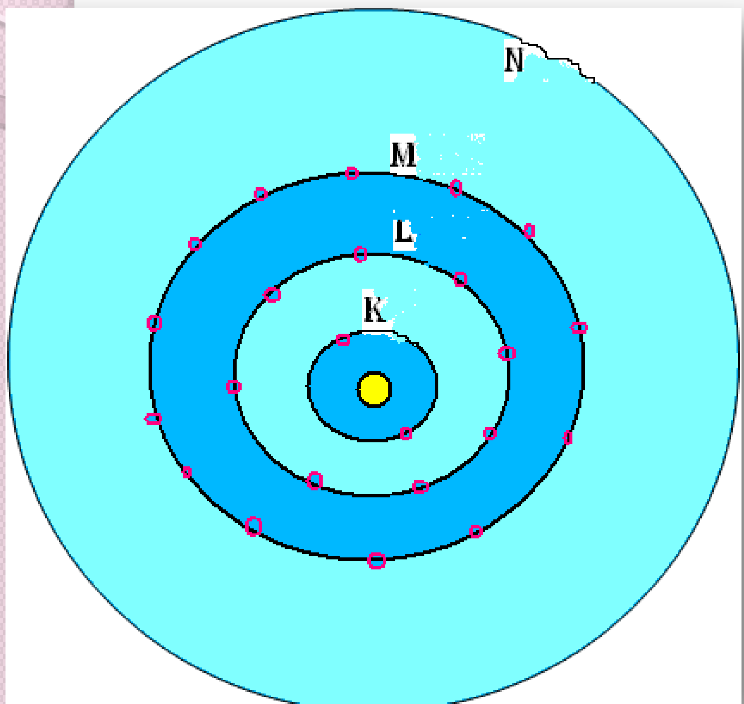
Postulates Bohr's model



- The energy change does not take place in a continuous manner.
- The frequency of radiation absorbed or emitted when transition occurs between two stationary states that differ in energy by ΔE, is given by

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

Postulates Bohr's model



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

- Where E_1 and E_2 are the energies of the lower and higher allowed energy states respectively.
- This expression is commonly known as Bohr's frequency rule.

Postulates Bohr's model

- The angular momentum of an electron in a given stationary state can be expressed as in equation

$$m_e vr = n.\frac{h}{2\pi} \quad n = 1,2,3.....$$

Thus an electron can move only in those orbits for which its angular momentum is integral multiple of $h/2\pi$ that is why only certain fixed orbits are allowed.

Limitations of Bohr's model

- It fails to account for the finer details (doublet, that is two closely spaced lines) of the hydrogen atom spectrum observed by using sophisticated spectroscopic techniques.
- This model is also unable to explain the spectrum of atoms other than hydrogen, for example, helium atom which possesses only two electrons.

Limitations of Bohr's model

- Further, Bohr's theory was also unable to explain the splitting of spectral lines in the presence of magnetic field (Zeeman effect)
- or an electric field (Stark effect).

- It could not explain the ability of atoms to form molecules by chemical bonds.