Department of Mathematics and Natural Sciences

CHE 101: Introduction to Chemistry

Lecturer: **Muhammad Mahfuz Hasan**

Lecture-02

Content: **1. Thomson Atomic Model**
2. Rutherford's atomic theory,
3. Bohr's atomic theory,
4. Spectrum

Discovery of the Electron

In 1897 the British physicist **J. J. Thomson** conducted a series of experiments that showed the atoms were not indivisible particles.

In this experiment, two electrodes from a high-voltage source are sealed into a glass tube from which the air has been evacuated(Low). The negative electrode is called cathode and the positive one is anode. When high-voltage current is turned on, the glass tube emits a greenish light. This greenish light is caused by interaction of the glass with cathode rays, which are rays that originate from the cathode.

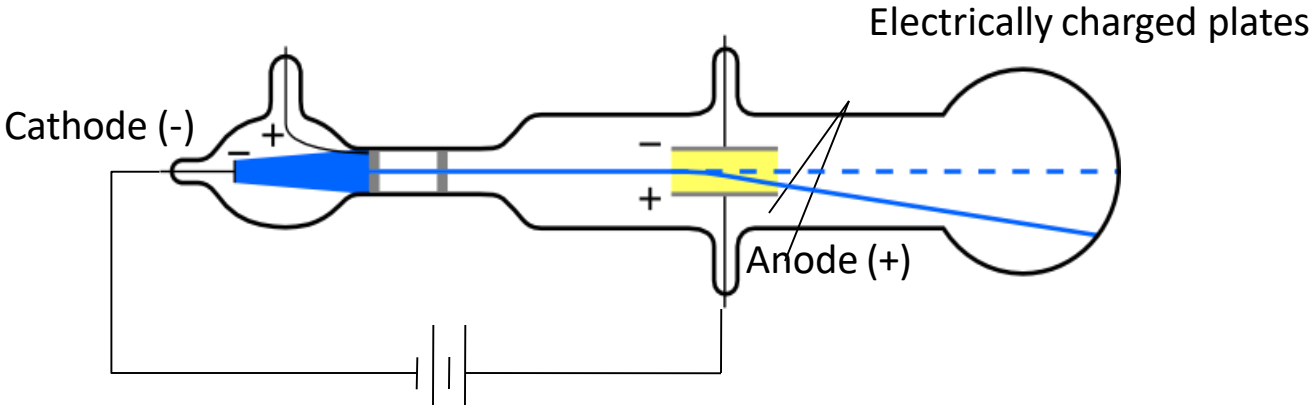
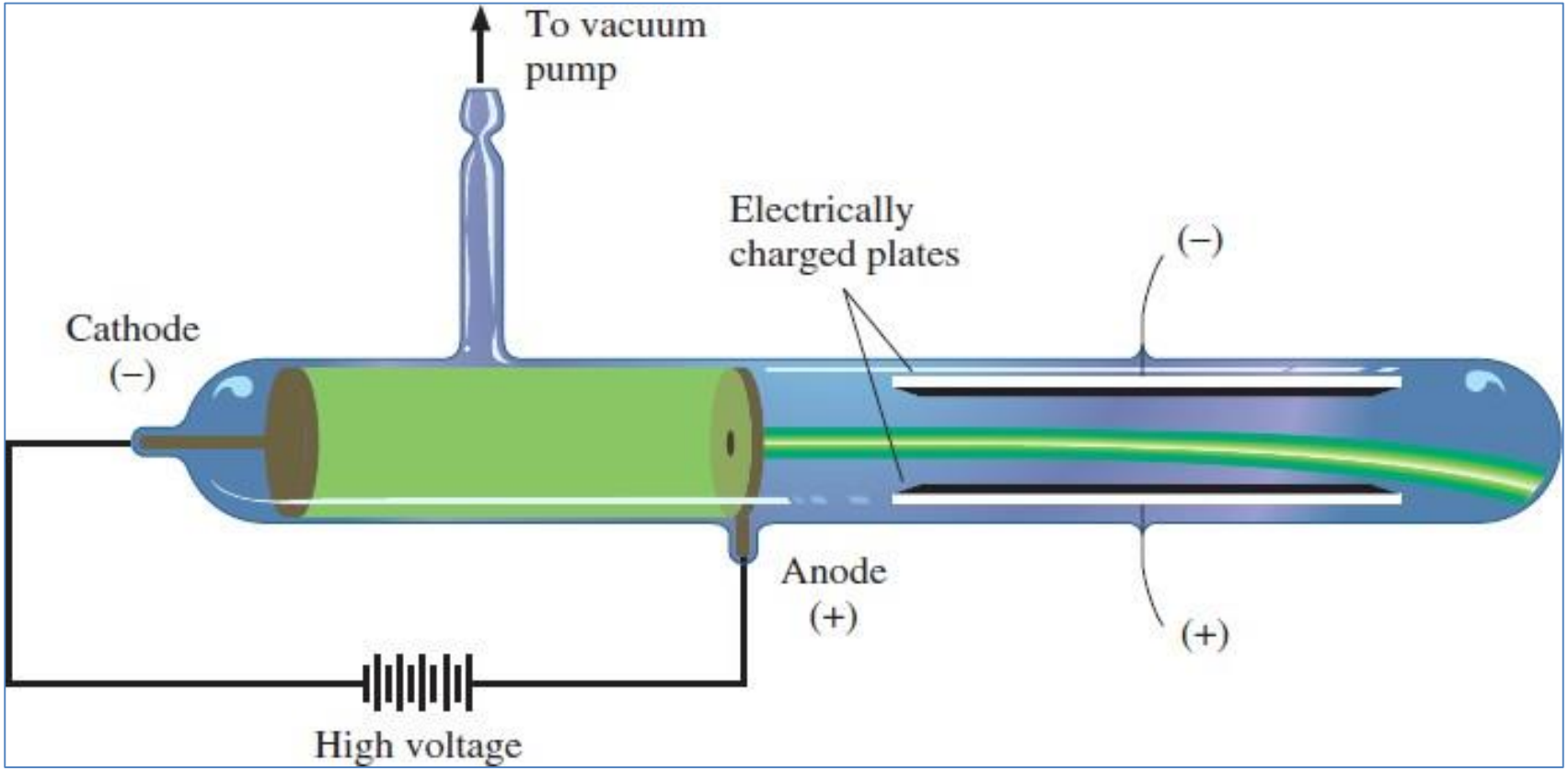
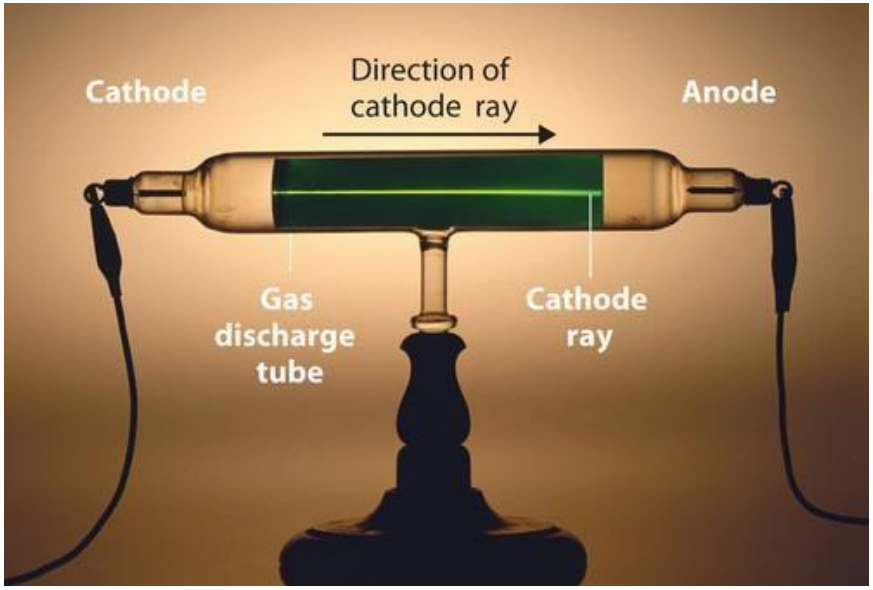


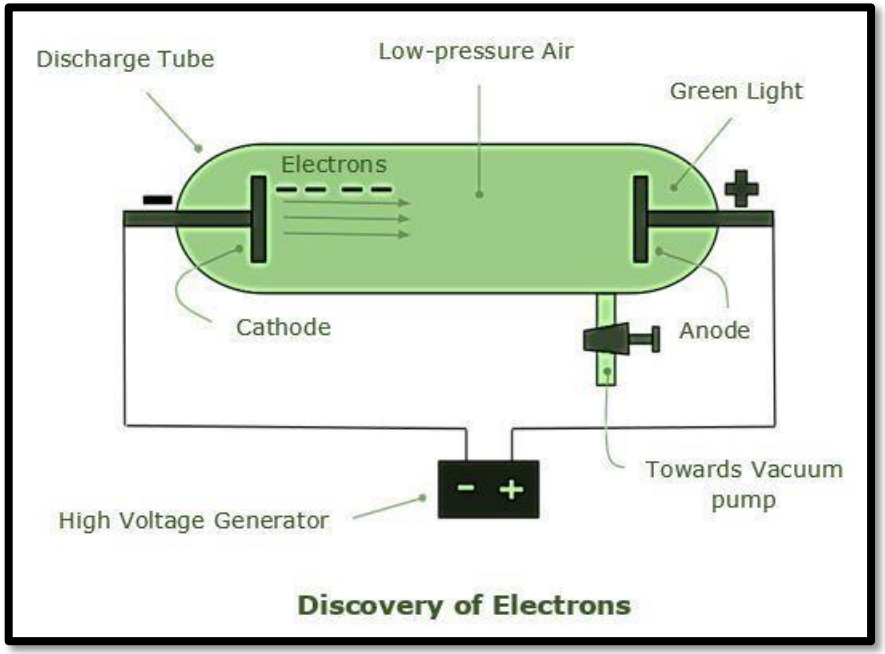
Fig: Formation of cathode rays

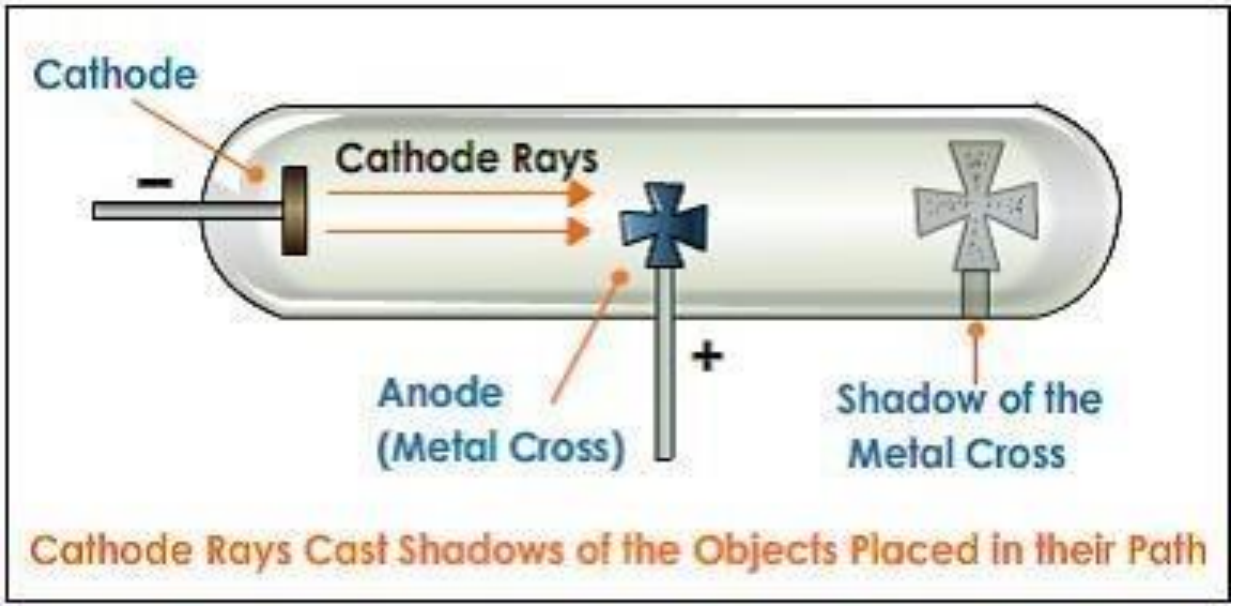
After the cathode rays leave the negative electrode, they move toward the anode where some rays pass through a hole from a beam. This beam bends away from the negatively charged plate and toward the positively charged plate. From such evidence Thomson concluded that a cathode ray consists of a beam of negatively charged particles (electrons) and that electrons are constituents of all matter.

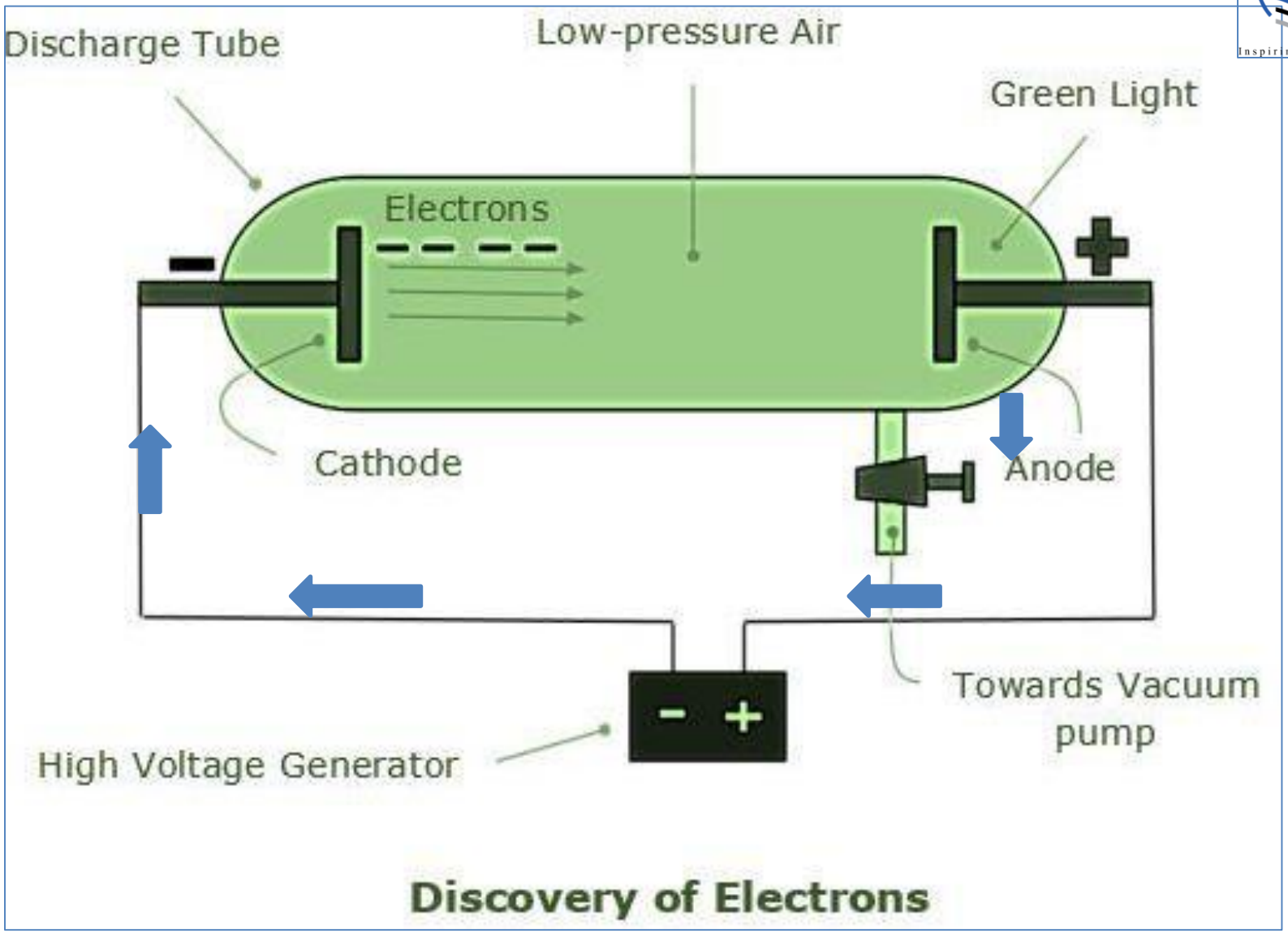


JJ Thomson E/M Experiment


Cathode
Direction of cathode ray
Anode
Gas discharge tube
Cathode ray

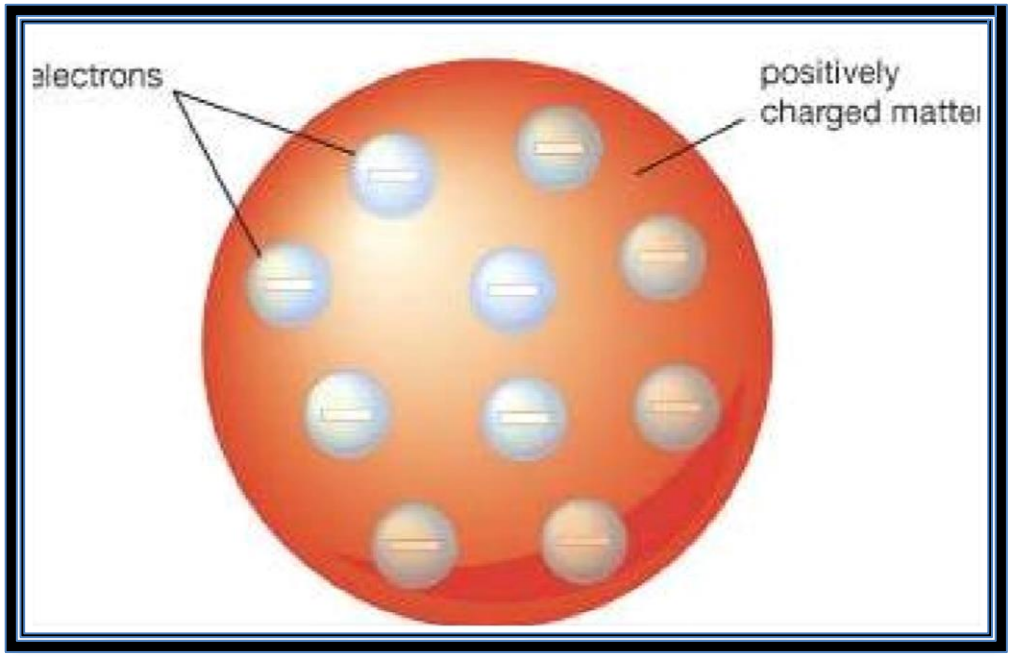

Discharge Tube
Low-pressure Air
Green Light
Electrons
Cathode
Anode
Towards Vacuum pump
High Voltage Generator
Discovery of Electrons

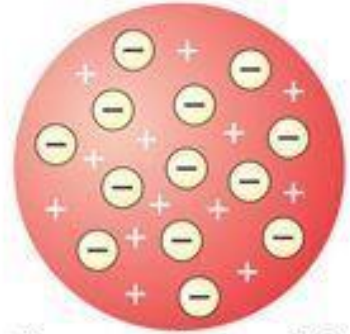

Cathode
Cathode Rays
Anode (Metal Cross)
Shadow of the Metal Cross
Cathode Rays Cast Shadows of the Objects Placed in their Path


Discharge Tube
Low-pressure Air
Green Light
Electrons
Cathode
Anode
Towards Vacuum pump
High Voltage Generator
Discovery of Electrons

Thomson Atomic Model

Between 1903 and 1907 Thomson tried to solve the mystery by adapting an atomic model that had been first proposed by Scottish scientist William Thomson (Lord Kelvin) in 1902. According to the Thomson atomic model, often referred to as the "plum-pudding" model, the atom is a sphere of uniformly distributed positive charge about one angstrom in diameter. Electrons are embedded/join in a regular pattern, like raisins in a plum pudding, to neutralize the positive charge.

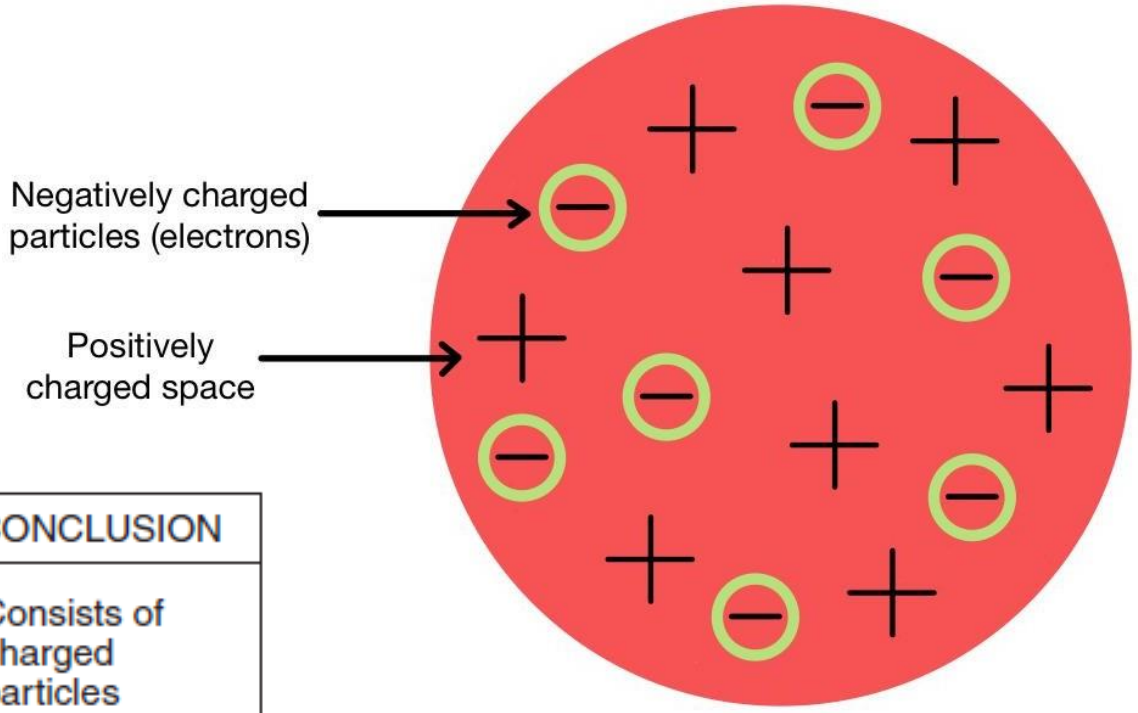




Thomson's model of an Atom

Plum pudding

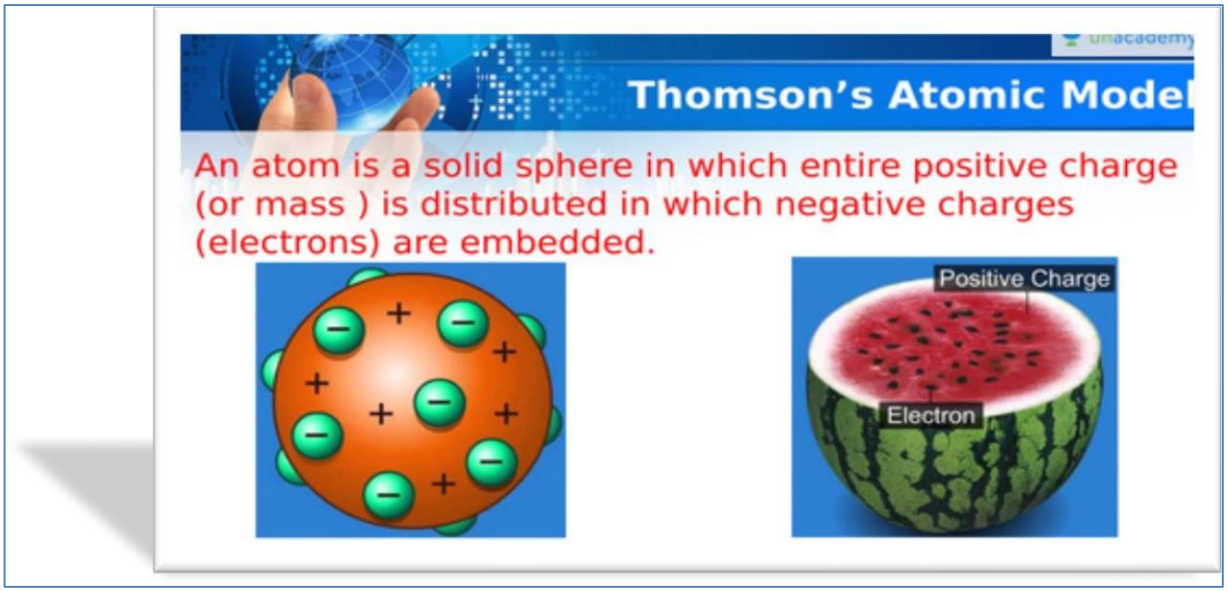
Plum Pudding Model

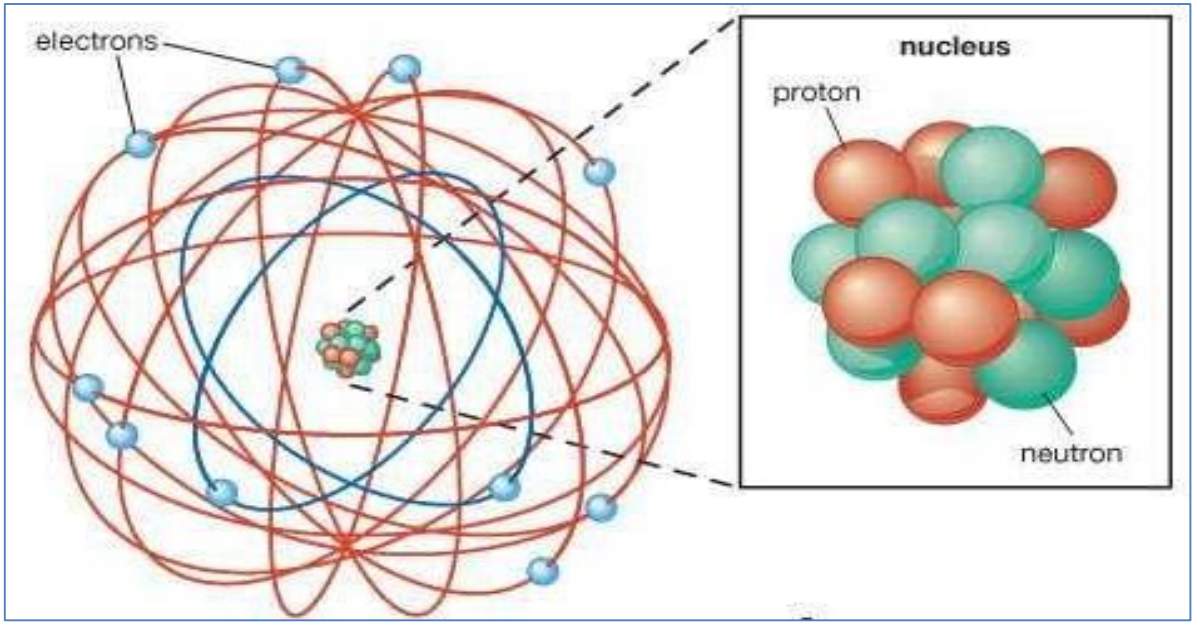


OBSERVATION	CONCLUSION
1. Ray bends in magnetic field.	Consists of charged particles
2. Ray bends toward positive plate in electric field.	Consists of negative particles
3. Ray is identical for any cathode.	Particles found in all matter

The **Plum Pudding Model** was a model of an atom created by J.J Thompson. The model describes the atom as negatively charged particles swimming in a positively charged sea. It got it's name because the model resembles the look of plum pudding

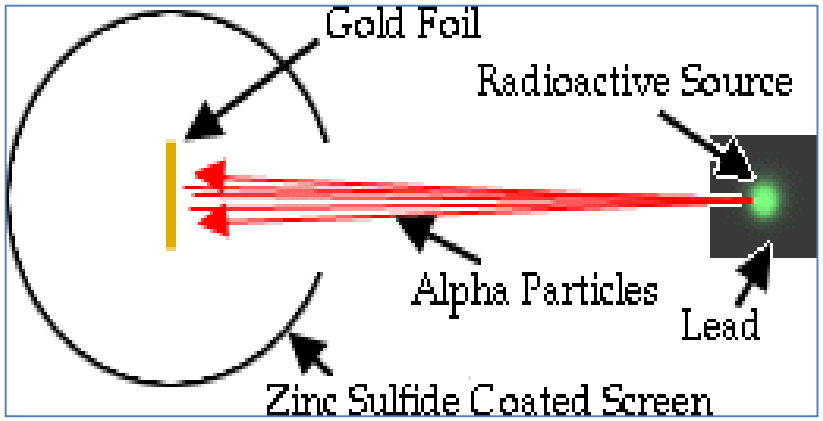
Thomson Atomic Model

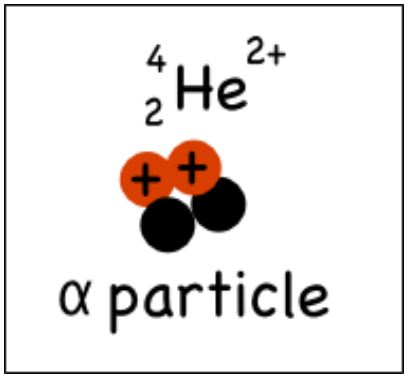




Idea of nucleus:
The gold foil experiment or α- scattering experiment

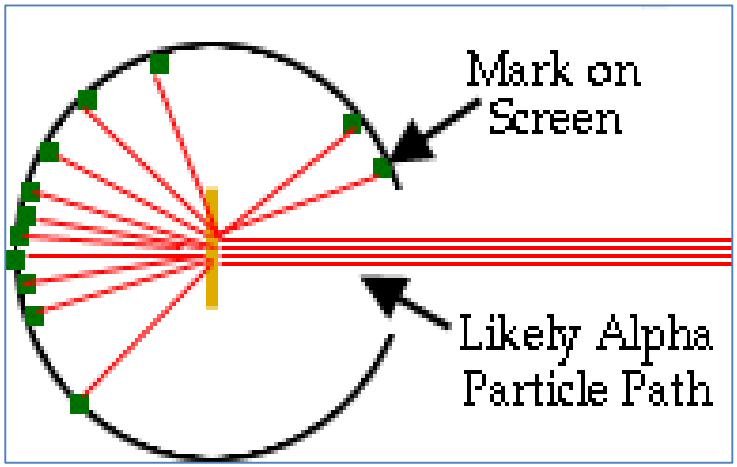
Rutherford in 1911 projected a beam of alpha (α) particles from a radioactive source upon a very thin gold foil. The α- particles emitted from radioactive elements with great velocities, on the average about 180,000 miles per second.

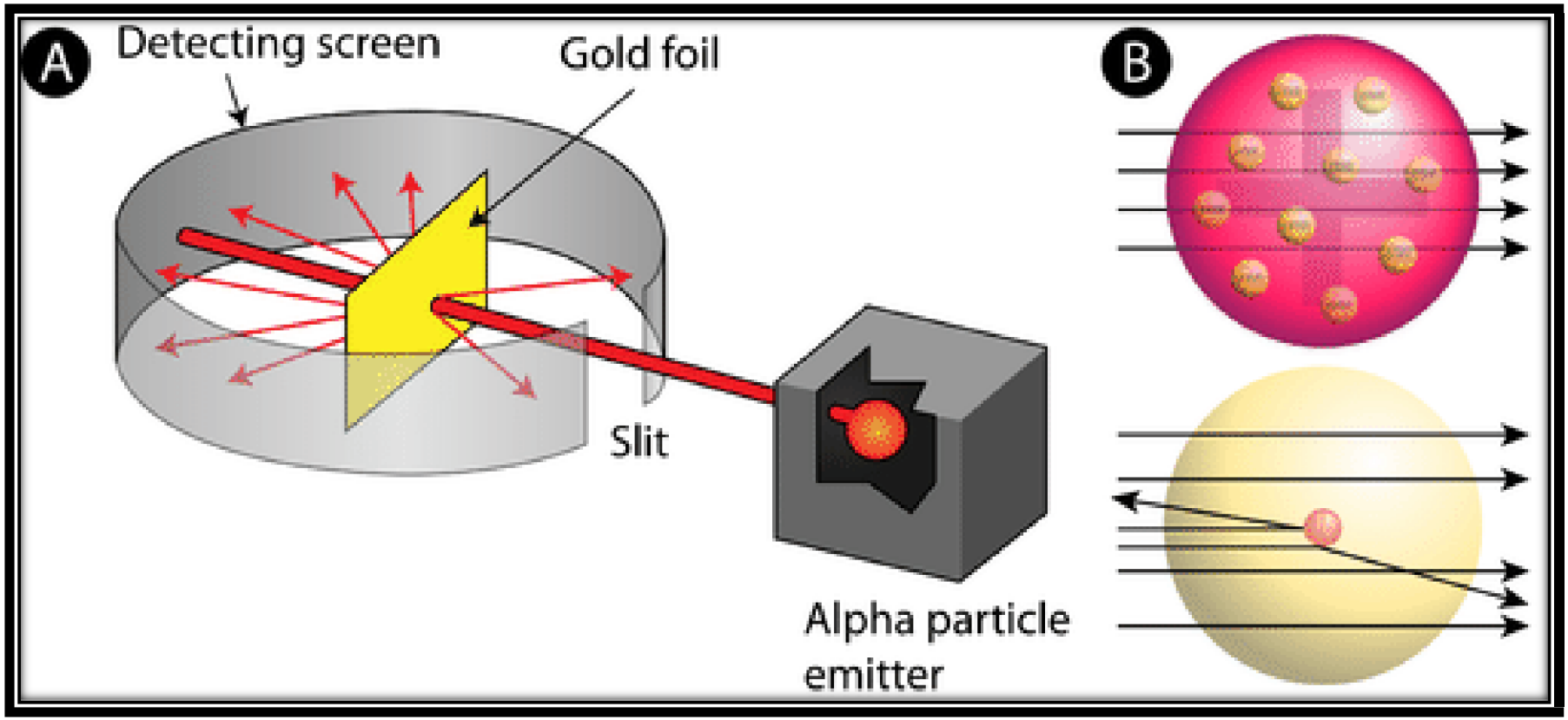




Observed results

1. most alpha particles were observed to pass straight through the gold foil without deflection
2. a few were scattered at large angles
3. some even bounced back toward the source as if the α-particles have met with some obstacles in their onward journey



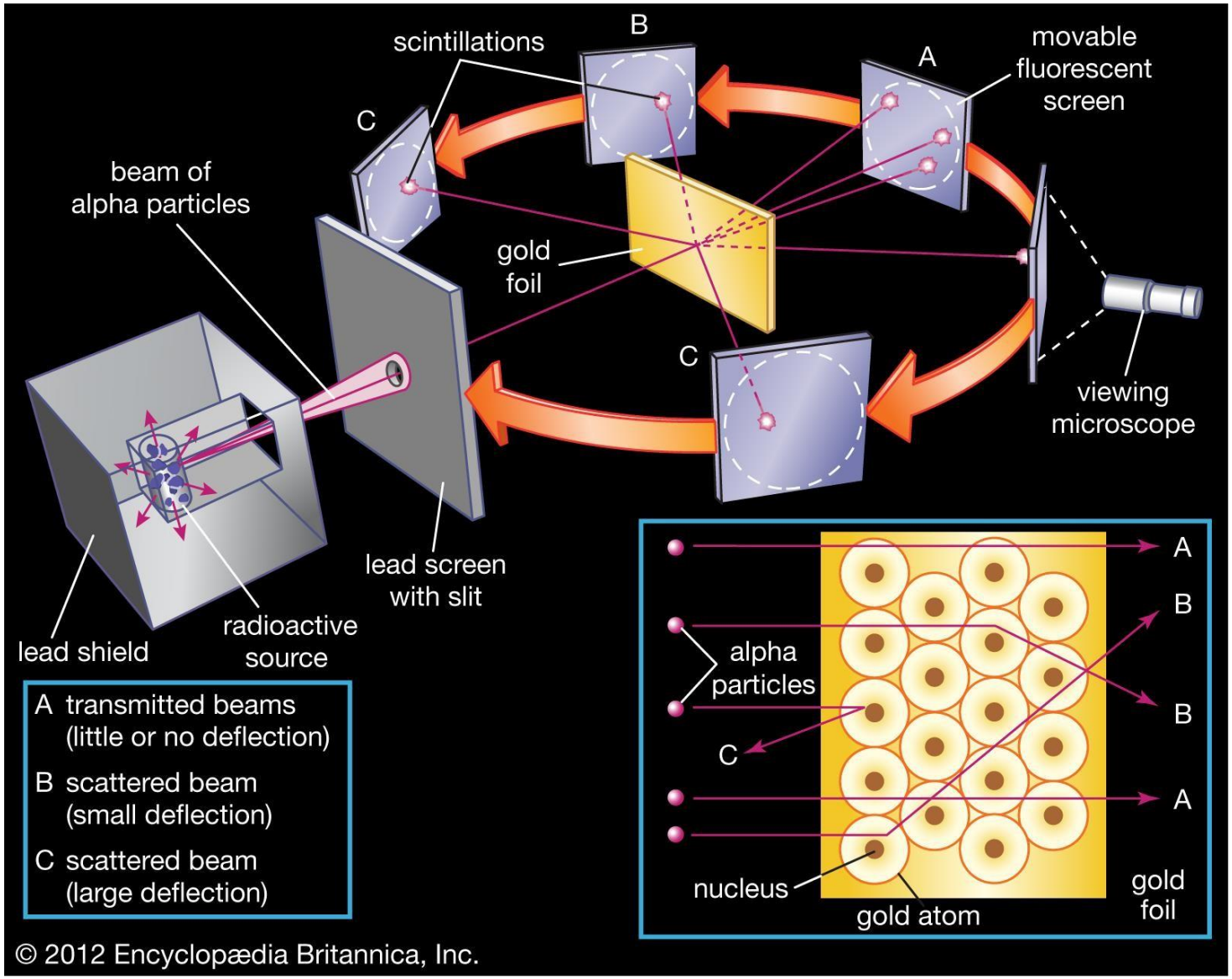


Gold Foil Experiment Rutherford

or

Alpha Particle Scattering Experiment

Rutherford's Atomic Model



Inferences/ Rutherford's Atom Model

(1) Most of the space of an atom is empty.

(2)Most of the mass of an atom is concentrated at the centre (99.95 % or more) – called the nucleus which is positively charged and exceedingly small as compared to the total size of atom.

(3)Electrons move around the nucleus, almost like the solar system in which the planets move around the sun.

(4) The number of electrons must be equal to the number of positively charged particles in the nucleus so that the atom as a whole is neutral.

(5) Due to rapid rotation of electron around the nucleus, the inward force of electrostatic attraction (centripetal force) between electron and nucleus is exactly counterbalanced by the outward centrifugal force.

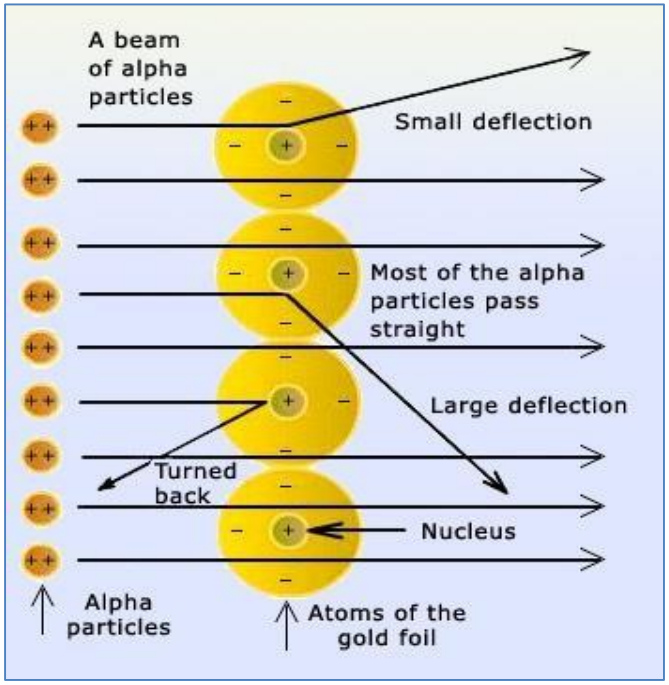


Fig: Deflection of α- particles by nuclei in a metal foil

❖The Rutherford atomicthe planetary model of the atom or Solar system model. model has been alternatively called the nuclear atom, or

❖Although most of the mass of an atom is in its nucleus, the nucleus occupies only a very small portion of the space of the atom. **Nuclei** have **diameters** of about 10^{-15} **m** whereas **atomic diameter** is about 10^{-10} **m**, a **hundred thousand times larger**.

Limitations of Rutherford Atom Model

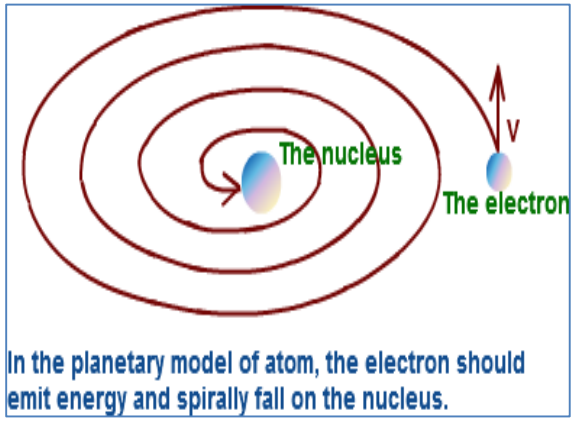
1. This model is based upon Newton's laws of motion and gravitation. But Newton's laws of motion and gravitation can only be applied to neutral bodies such as planets and not to charged bodies such as tiny electrons moving round a positive nucleus. The analogy does not hold good since the electrons in an atom repel one another, whereas planets attract each other because of gravitational forces. Besides there is electrostatic attraction in a nuclear atom model.

2. According to Maxwell's electromagnetic theory, any charged body such as electrons rotating in an orbit must radiate energy continuously thereby losing kinetic energy. Thus rotating electron will lose energy and its orbit will become smaller and smaller and it will ultimately fall into the nucleus following a spiral path, annihilating the atom model.

3. If there is continuous emission of radiation, the spectra of an atom will be a band or continuous spectra. But an atom gives discontinuous or line spectra.

4. Rutherford did not give any idea about the shape of the orbits.

5. There was no explanation about the rotation of electrons in an atom with many electrons.

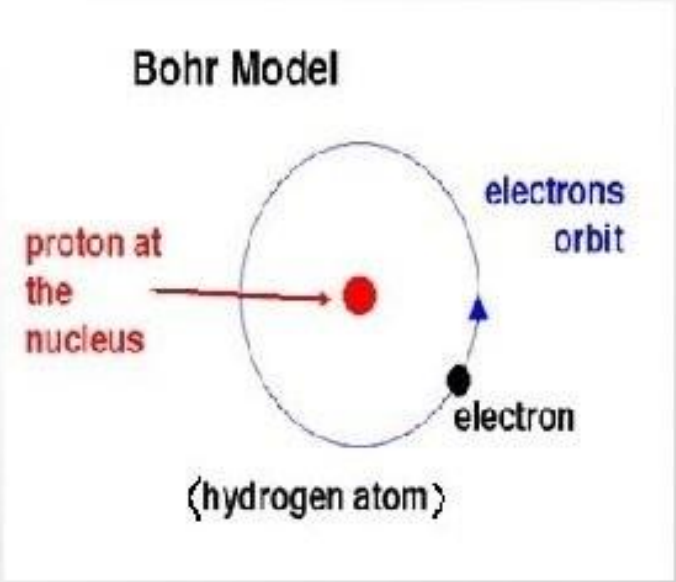


In the planetary model of atom, the electron should emit energy and spirally fall on the nucleus.

Niels Bohr

(1885 - 1962)

- The **Bohr Model**
- Suggests that electrons travel around the nucleus in orbits



Bohr's Atom Model

1. Postulates of Energy Levels

- An atom has a number of stable orbits in which an electron can revolve without the radiation of energy. These orbits are referred to as 'Energy Levels'(Stationary orbit).

- An electron moving in an orbit can have only certain amount of energy, not an infinite number of values, i.e., its energy is quantized.
- While rotating in an orbit an electron does not absorb or emit energy. Radiation is absorbed or emitted when an electron moves from one orbit to another.
- The energy that an electron needs in order to move in a particular orbit depends on the radius of the orbit. An electron in an orbit distant from the nucleus requires higher energy than an electron in an orbit near the nucleus.(The lowest energy is found in the smallest orbit)
- If the electrons moves in a circular orbit, its motion is subject to the ordinary laws of electrical and centrifugal force. These orbits are decided by the condition that the angular momentum of the electron in such orbit must be an integral multiple of $h/2\pi$, that is,

$$mvr = nh/2\pi$$

where, m is the mass of the electron, v is its velocity, r is the radius of the orbit, h is the Plank's constant and n is 1, 2, 3, 4, 5 etc.

2. Postulates of radiation of energy

When an electron jumps from one energy level to another there is absorption or emission of energy. When an electron moves from a higher energy level to lower energy level there is emission of radiation and when electron moves from a lower level to a higher one there is absorption of radiation. The energy radiated is electromagnetic.

If E_1 and E_2 are energies of electron in the initial and final levels respectively, the difference of energy radiated when the electron passes from the higher to the lower energy level is given by the relation: $E_2 - E_1 = h\upsilon$
where h is the Plank's constant and υ is the frequency of radiation.

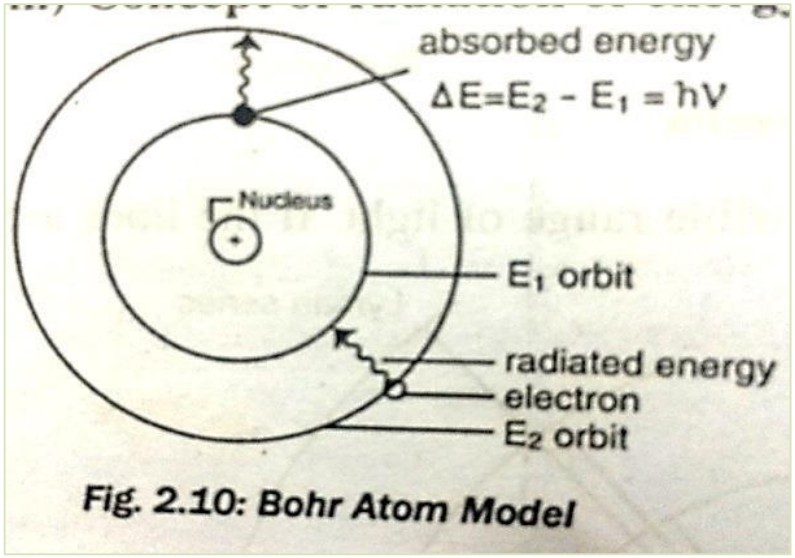
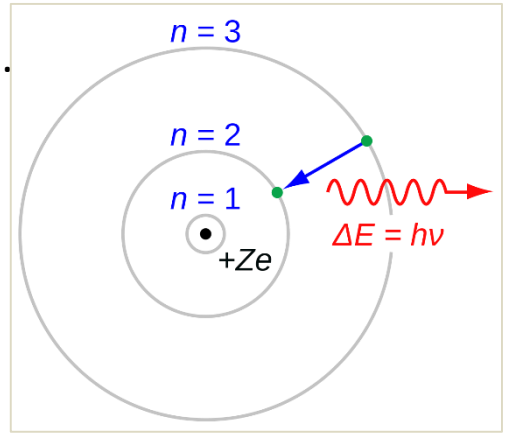


Fig. 2.10: Bohr Atom Model



+Ze indicates the **positive nuclear charge** where Z is an integer and Z=1,2,3……. And Z indicates the atomic number or proton numbers which is equal to electrons

Since these transitions result in the emission of a photon of frequency ν and energy h ν, we can write-

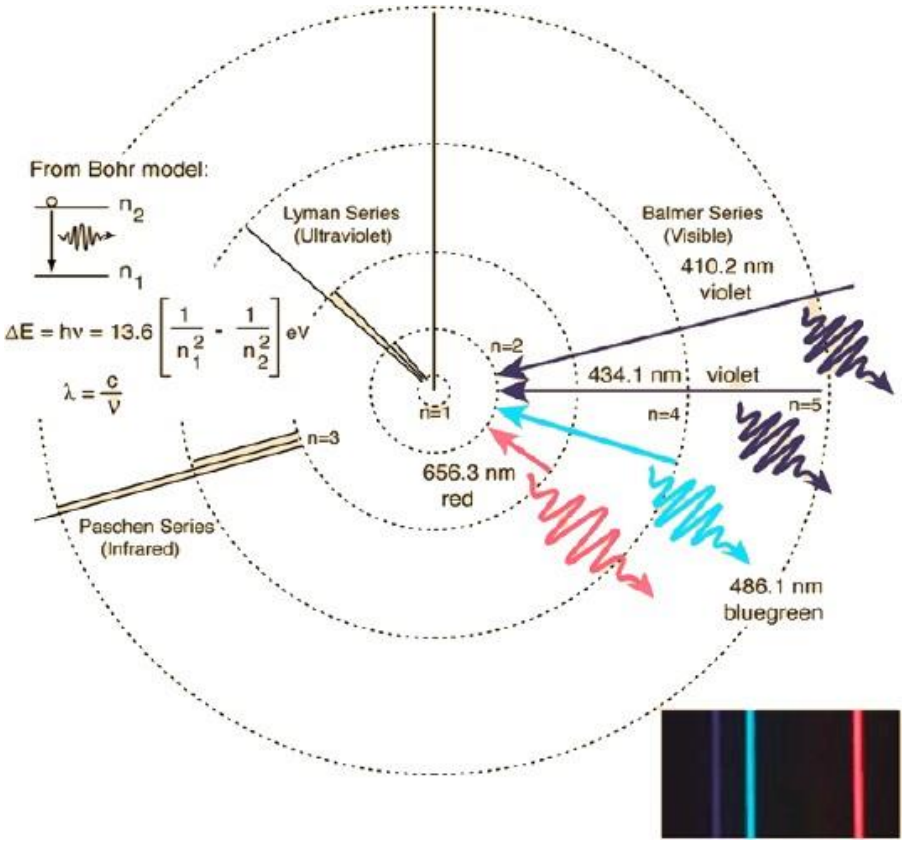
$$\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$ where $n_i > n_f$

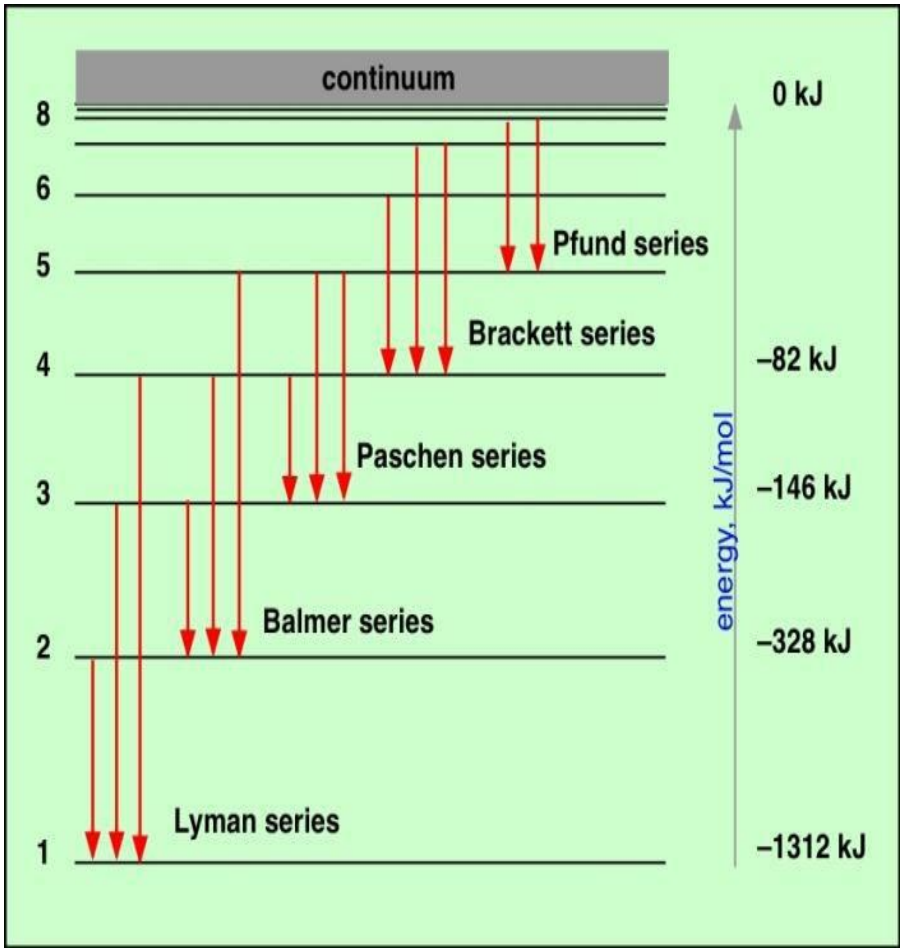
$R = R_H$ = Rydberg constant = 109737 cm^{-1}

Rydberg equation: $$\overline{\nu} = \frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{ cm}^{-1}$$

Series	n_1	n_2	Region of spectrum
Lyman	1	2,3,4…..	Ultraviolet
Balmer	2	3,4,5…..	Visible
Paschen	3	4,5,6…..	Infrared
Bracket	4	5,6,7…..	Infrared
Pfund	5	6,7,8…..	Infrared

H-atom spectrum





In 1914 Lymann

$$\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{n^2} \right) \text{ m}^{-1}, \text{ n = 2, 3, 4, ---}$$

In 1885 J. J. Balmer

Balmer expressed the wavelengths of these lines by an equation-

$$\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \text{ m}^{-1}, \text{ n = 3, 4, 5, ---}$$

In 1908 Paschen

$$\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{3^2} - \frac{1}{n^2} \right) \text{ m}^{-1}, \text{ n = 4, 5, 6, ---}$$

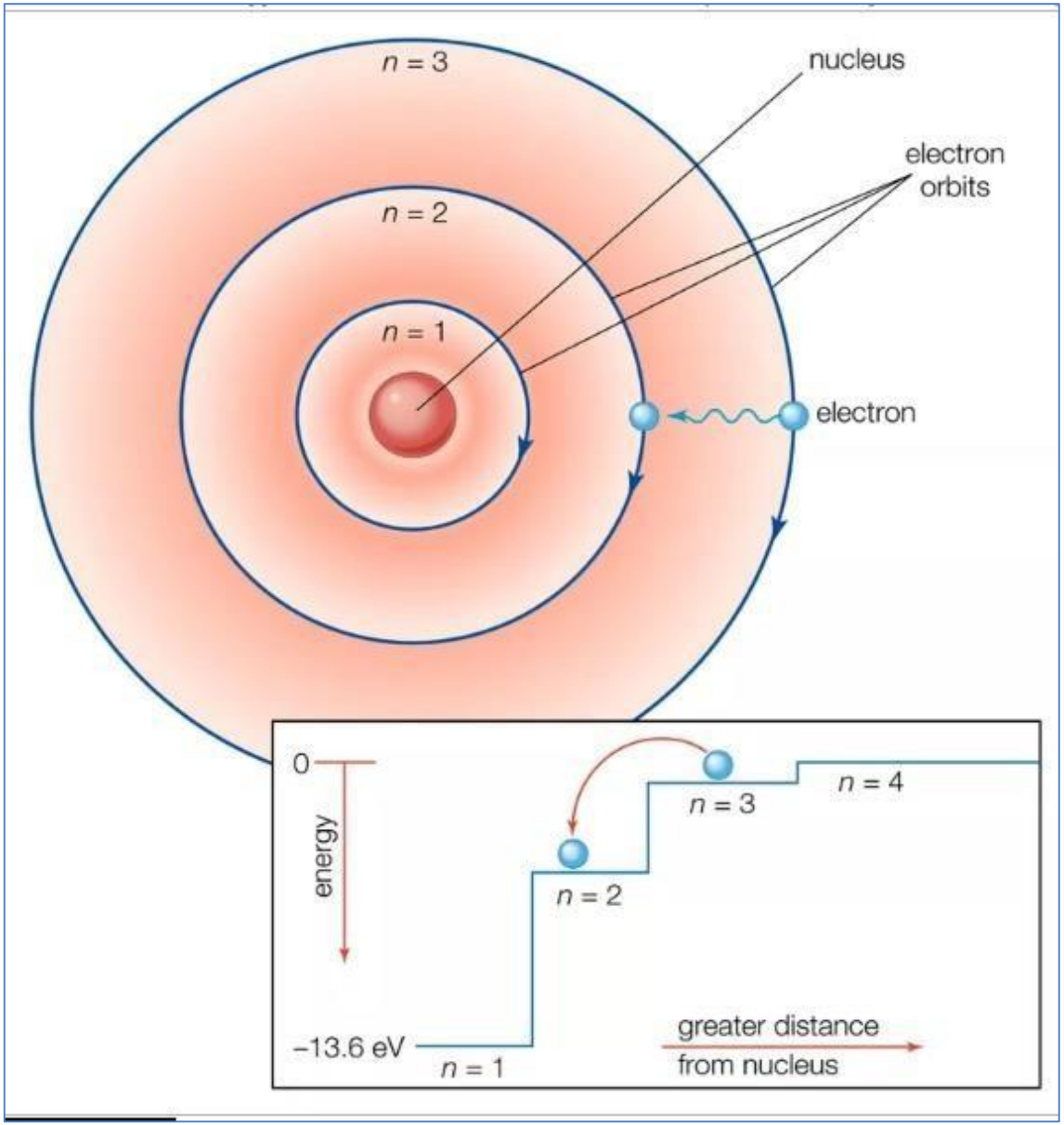
In 1922 Brackett

$$\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{4^2} - \frac{1}{n^2} \right) \text{ m}^{-1}, \text{ n = 5, 6, 7, ---}$$

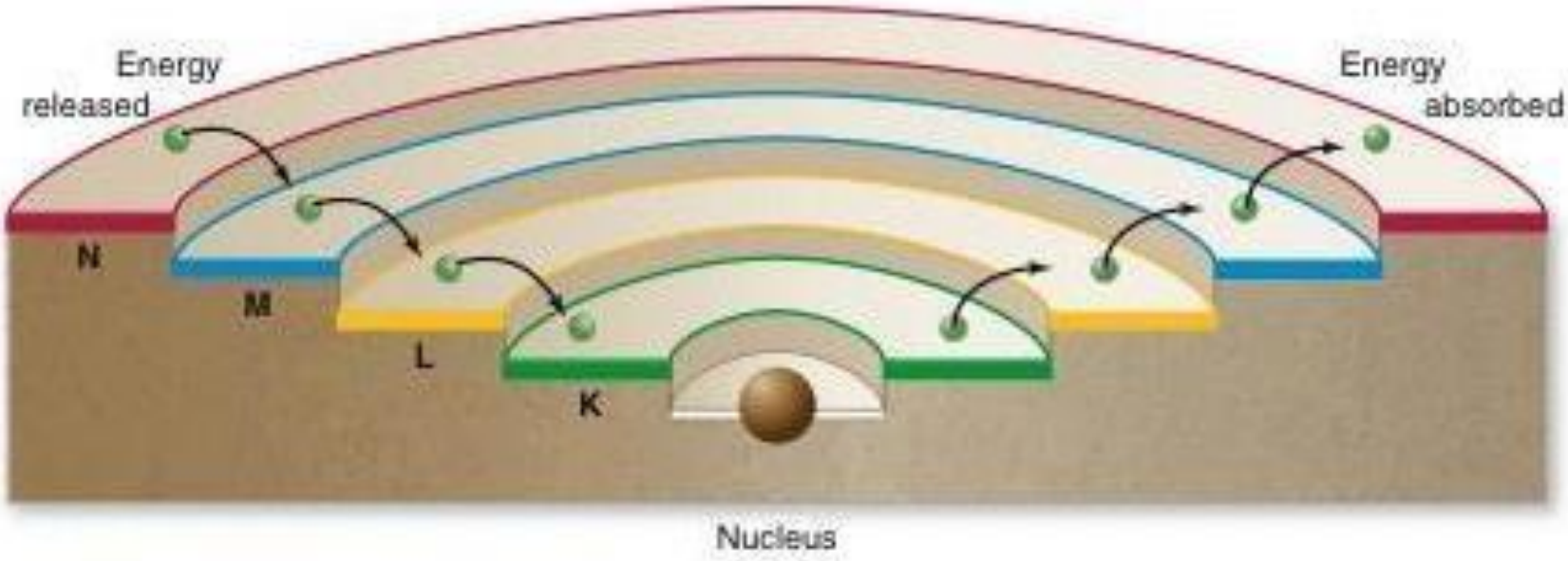
In 1924 Pfund

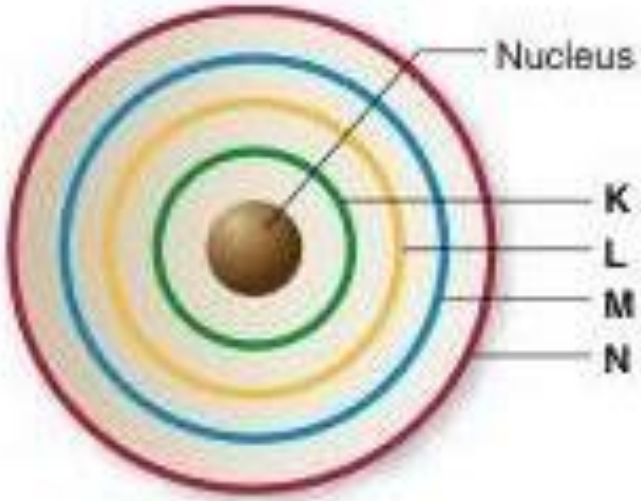
$$\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{5^2} - \frac{1}{n^2} \right) \text{ m}^{-1}, \text{ n = 6, 7, 8,---}$$

Bohr's Atomic Model

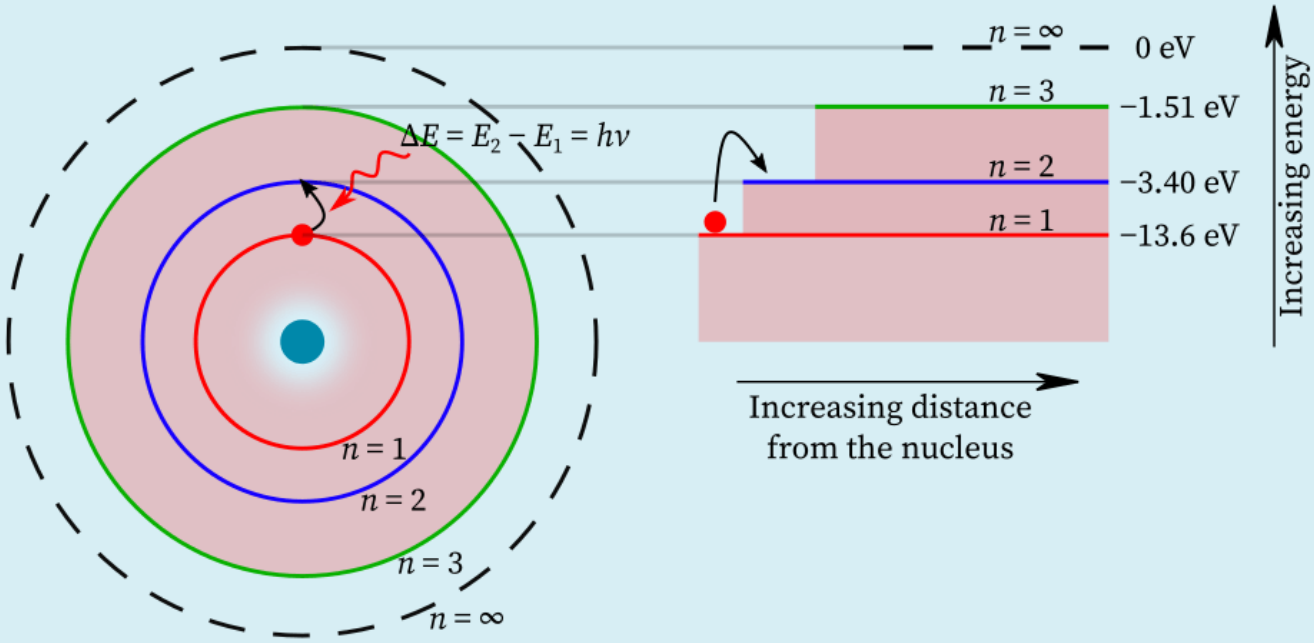


ENERGY LEVELS





Bohr's Atomic Model

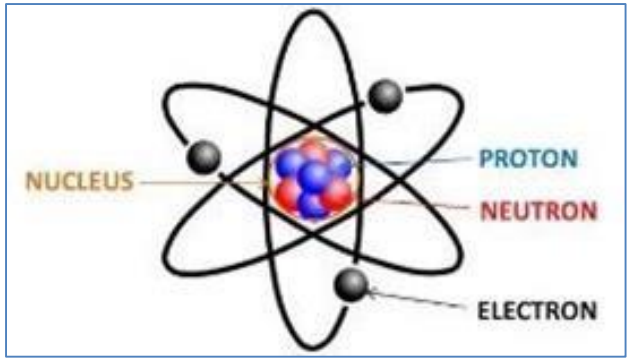


Limitations of Bohr Theory

1. Bohr theory successfully explain spectrum of hydrogen, but can not explain **the spectral lines of atoms with more than one electron.**

2. This theory is unrealistic in the sense that periodic motion around a central body usually follows an **elliptical path rather than a circular path (Fig)which** has been assumed in the case of Bohr theory. If electrons follow elliptical path, the velocity along the path does not remain constant.

3.According to Bohr's model when an electron jumps from one energy level to another, a single line is supposed to appear on the spectra. However, when a **spectrograph is developed with high resolving power there are two or more lines very close together are observed**. Bohr's atom model gives no explanation on this.



Limitations of Bohr Theory

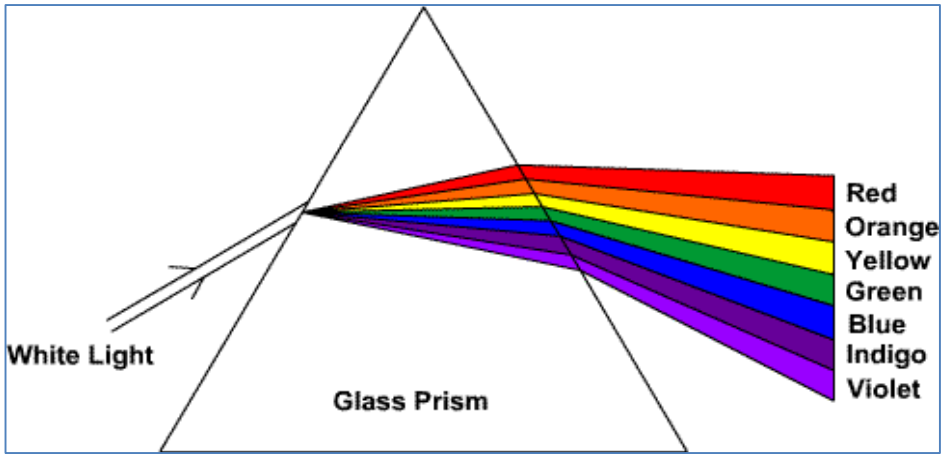
5. It violates the Heisenberg Uncertainty Principle. The Bohr atomic model theory considers electrons to have both a known radius and orbit i.e. known position and momentum at the same time, which is impossible according to Heisenberg.

6. The Bohr atomic model theory made correct predictions for smaller sized atoms like hydrogen, but poor spectral predictions are obtained when larger atoms are considered.

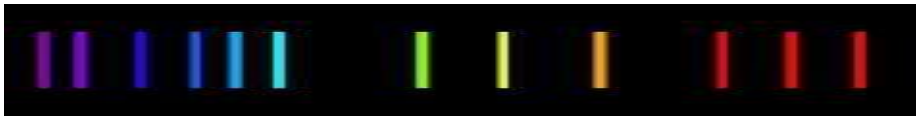
Spectrum and hydrogen spectrum

When white light is passed through a prism it is resolved into various color components of different wave length, and what is obtained is called a **spectrum.**

In the seventeenth century Newton showed that sun light is composed of various color components (which can be combined to produce white light as well).

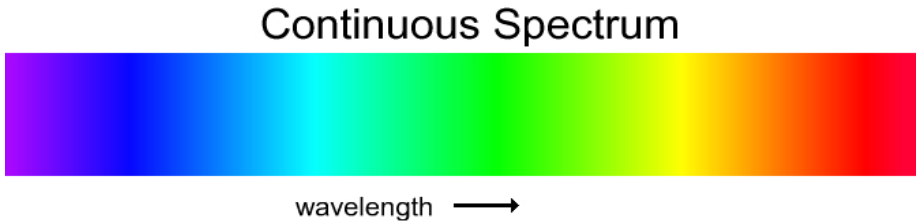


Line Spectrum: A spectrum showing only certain colors or specific wavelengths of light.



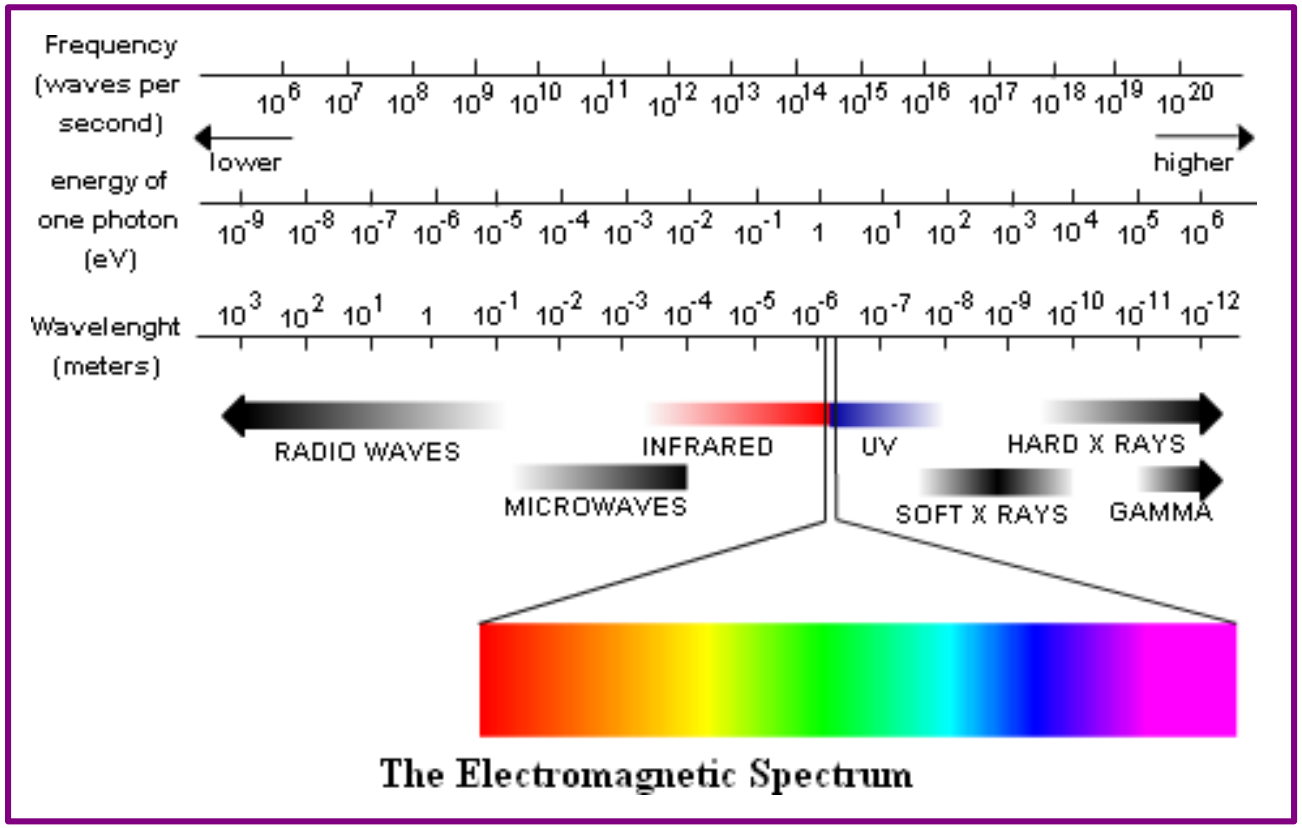
Continuous spectrum: A spectrum containing light of all wavelengths.

Example: rainbow



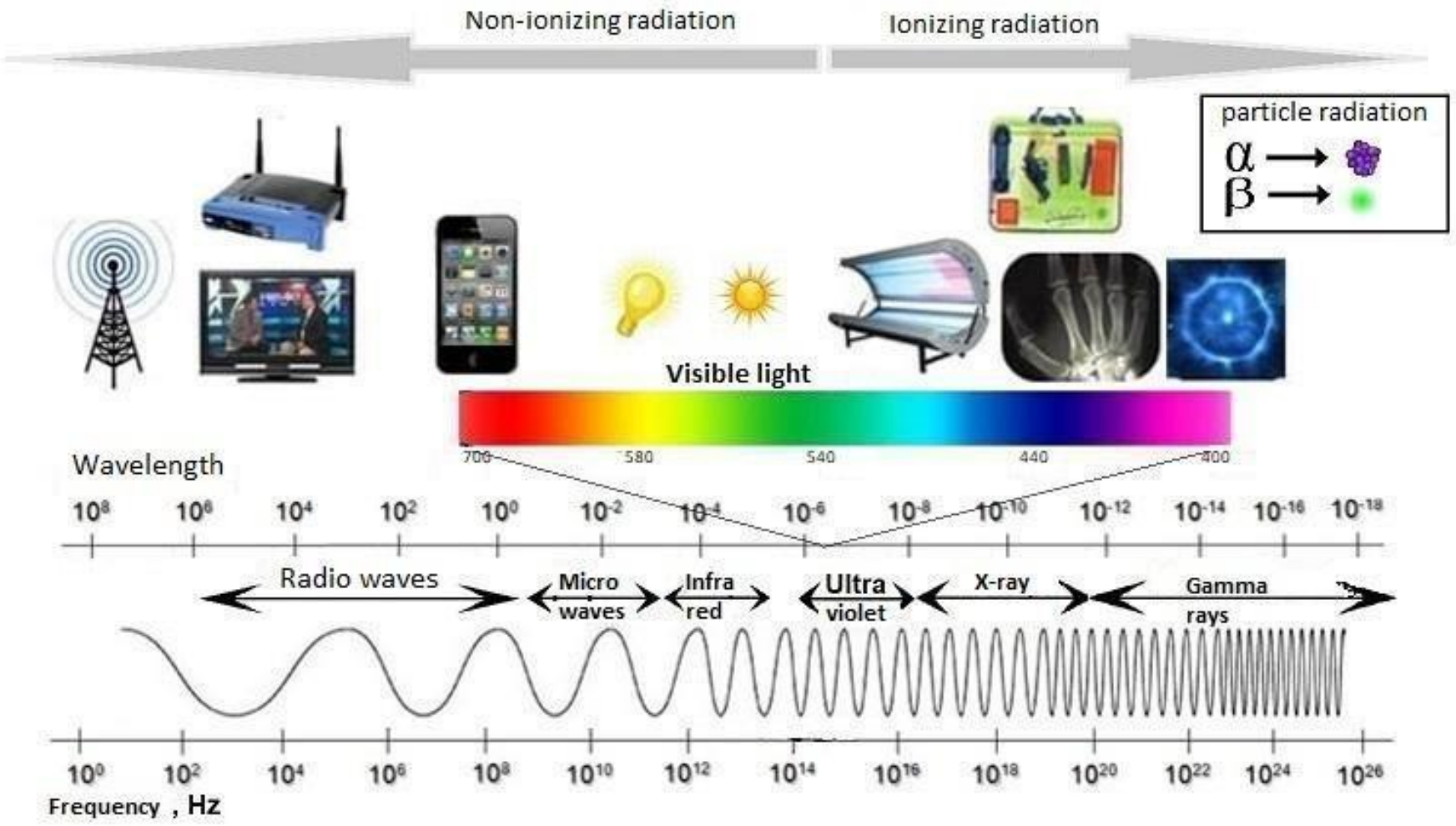
Nature and Origin of Radiation

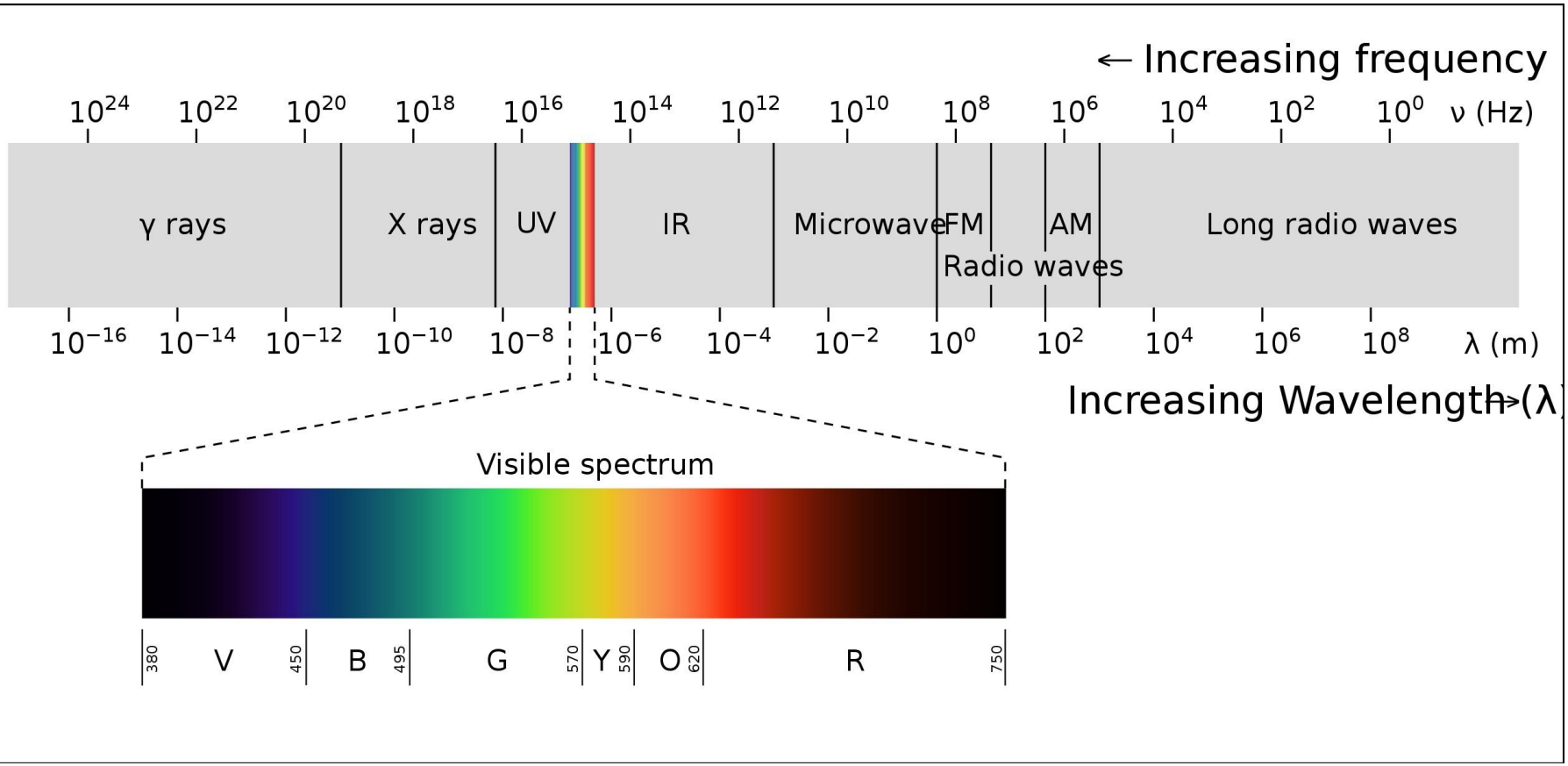
There are various kind of radiation which can be classified in electromagnetic radiation (EM) and particle radiation (p). The X-rays and γ-rays are part of the electromagnetic spectrum; both have a wavelength range between 10^{-4} and 10^{1} nm, they differ only in their origin.



The Electromagnetic Spectrum

The electromagnetic spectrum




← Increasing frequency
10^{24} 10^{22} 10^{20} 10^{18} 10^{16} 10^{14} 10^{12} 10^{10} 10^{8} 10^{6} 10^{4} 10^{2} 10^{0} ν (Hz)
γ rays
X rays
UV
IR
Microwave
FM
AM
Radio waves
Long radio waves
10^{-16} 10^{-14} 10^{-12} 10^{-10} 10^{-8} 10^{-6} 10^{-4} 10^{-2} 10^{0} 10^{2} 10^{4} 10^{6} 10^{8} λ (m)
Increasing Wavelength →(λ)
Visible spectrum
380 V 450 B 495 G 570 Y 590 O 620 R 750

Q: Is the spectrum for the incandescent light a continuous spectrum or a line spectrum?

Ans: Incandescent bulb is equivalent to white light, i.e it has all the seven wavelengths of light. Therefore when passed through a prism it is disperses into all the seven wavelengths without any spaces or boundaries and is obtained on the screen as a continuous spectrum.

Emission spectrum: If atoms or molecules are heated to sufficiently high temperature, they emit light of certain wavelengths. The pattern of frequencies emitted by the substance is called emission spectrum. From the emission spectrum we get bright lines.

Q: How is a bright line spectrum produced?

When electrons jump from the excited state to the ground state, the electrons emit energy in the form of light, producing a bright-line spectrum. Each element has its own unique bright-line spectrum.

Absorption spectrum: When white light is passed through a substance, black lines appear in the spectrum where light of some wavelengths have been absorbed by the substance. The pattern of frequencies absorbed by the substance is called absorption spectrum.

From the absorption spectrum we get dark lines.

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