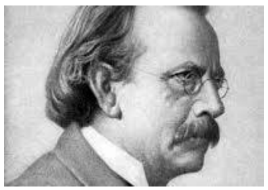
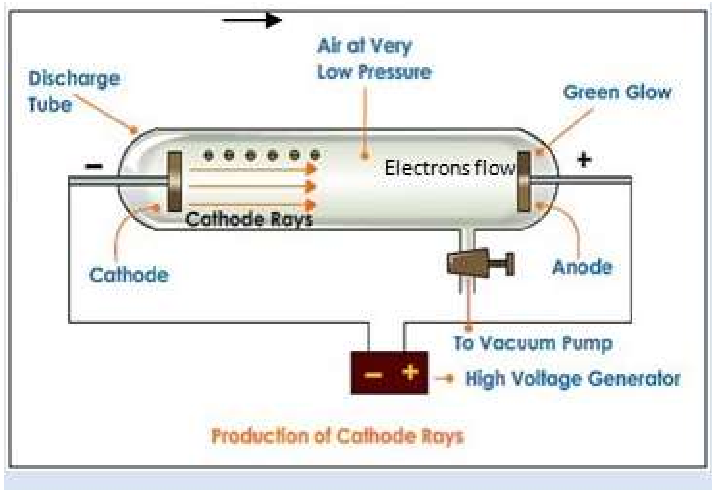
Atomic Models Revisit

The predecessors

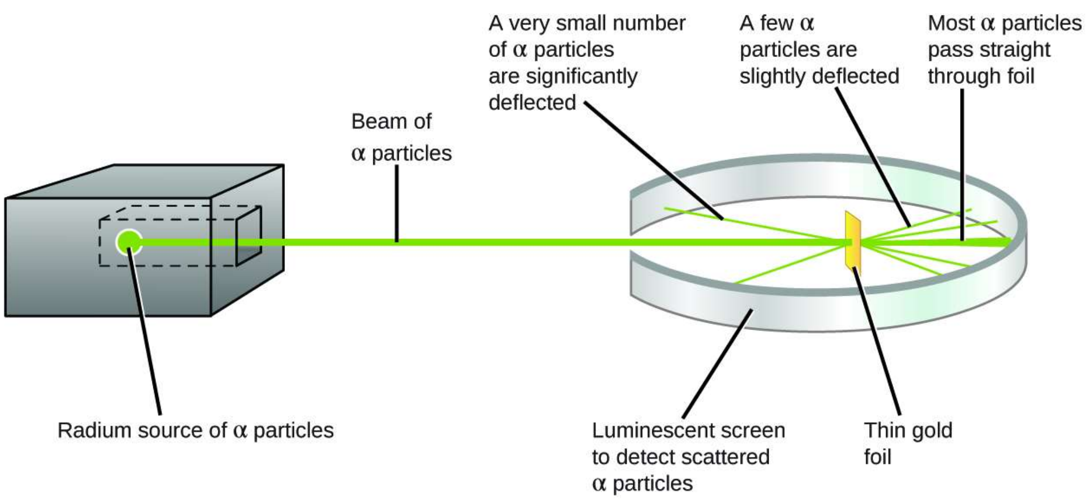


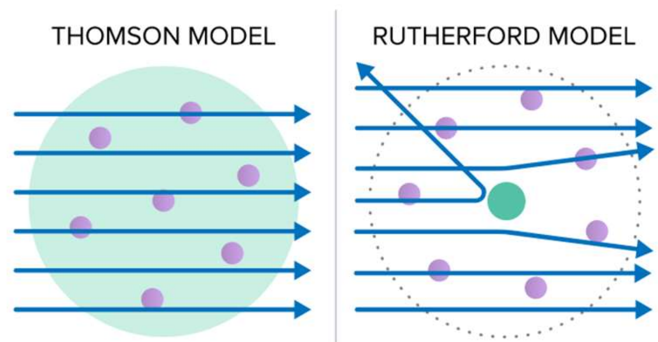
J. J. Thomson

Ernest Rutherford



A plum pudding



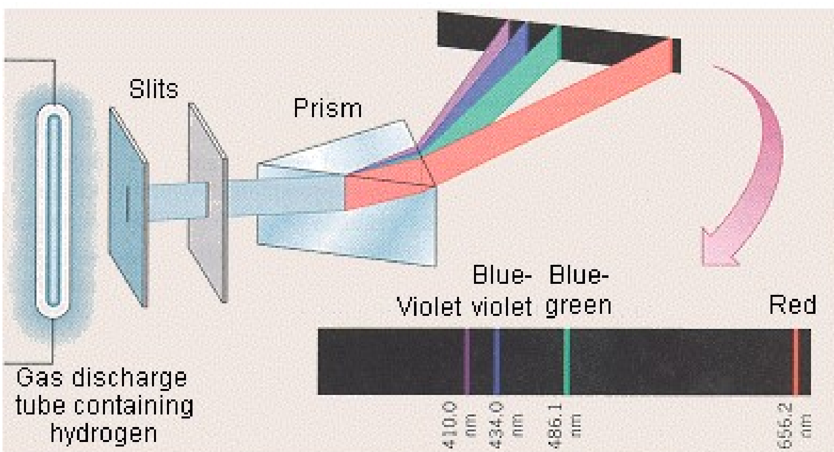


Rutherford's Model

- Positively charged – nucleus at the centre.
- Negatively charged – electrons revolves around the nucleus. Number of electrons = number of positive charge.
- Mass of nucleus = mass of protons + mass of neutrons ≈ mass of the atom.
- The nucleus is very small and the space surrounding the nucleus is mostly hollow.

Hydrogen spectra

Johann Jakob Balmer



Drawback Rutherford's model!!!

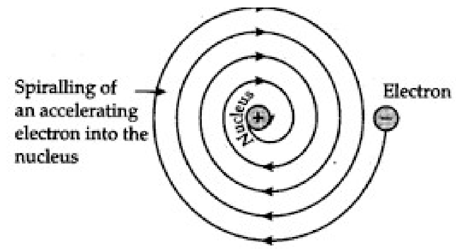
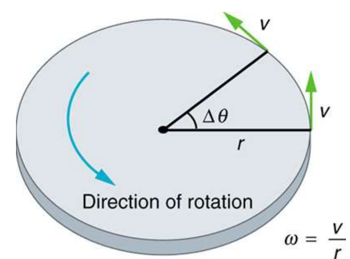


Fig. Spiralling of an accelerating electron into the nucleus with progressive loss of energy

Classical electrodynamics: *If a charged particle revolves around an oppositely charged particle it radiates energy continuously.*

$$E_{\text{rotational}} = \frac{1}{2} I\omega^2$$



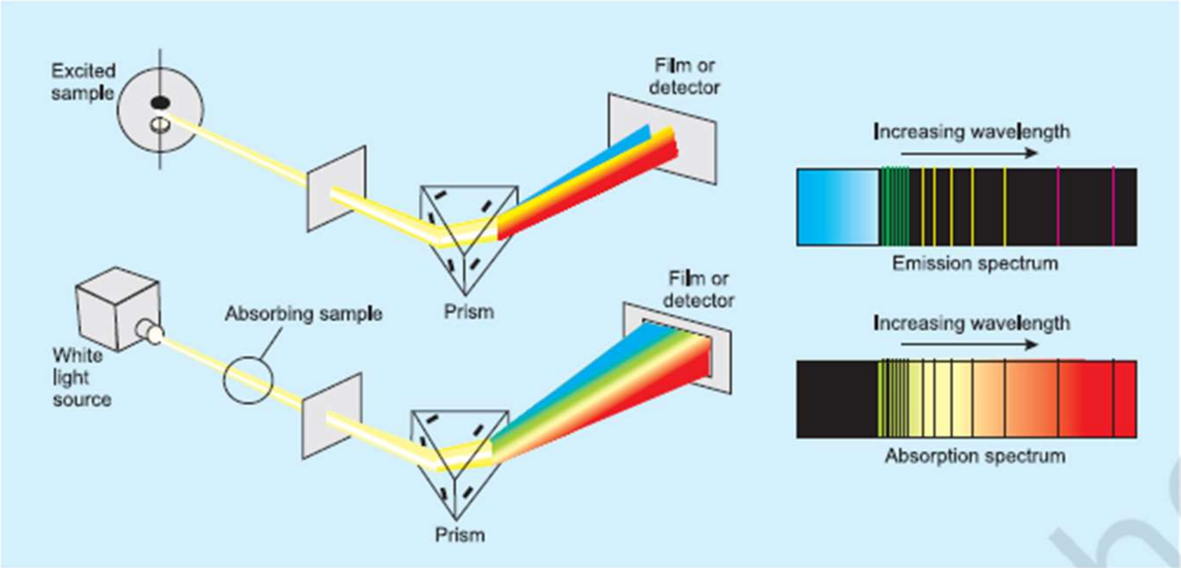
Continuous and Line spectra

- *The spectrum of white light, that we can see, ranges from violet at 7.50×10^{14} Hz to red at 4×10^{14} Hz. Such a spectrum is called **continuous spectrum**.*
- *The emission spectra of atoms in the gas phase, on the other hand, emit light only at specific wavelengths with dark spaces between them. Such spectra are called **line spectra** or **atomic spectra** because the emitted radiation is identified by the appearance of bright lines in the spectra.*

Emission and Absorption Spectra

(a) Atomic emission. The light emitted by a sample of excited atoms can be passed through a prism and separated into certain discrete wavelengths. Although a single atom can be in only one excited state at a time, the collection of atoms contains all possible excited states.

(b) Atomic absorption. When white light is passed through unexcited atoms and then through a slit and prism, the transmitted light is lacking in intensity at the same wavelengths as are emitted in (a).



Line Spectrum of Hydrogen

Balmer's observations

$$\bar{\nu} = 109677\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

n = 3, 4, 5…

Johannes Rydberg

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

n_1 = 1, 2 , 3, ….

n_2 = n_1+1, n_1+2, n_1+3, …

R_H = 109,677 cm^{-1} and is called the **Rydberg constant** for 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

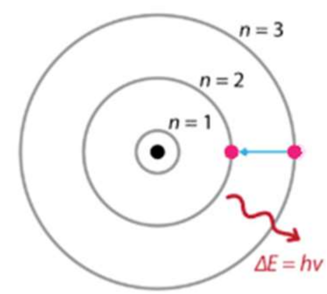
Bohr's Atomic Model

Niels Bohr

(a) The electron is bound in a circular orbit around the nucleus such that the angular momentum is quantized in integral units of Planck's constant,

$$\boldsymbol{mvr = \frac{nh}{2\pi}};\ \mathrm{m = mass\ of\ electron, v = velocity\ of\ electron, r = radius\ of\ the\ orbit}$$

The energy levels (E_r, r= 1, 2, 3, …) are represented as K, L, M, N etc. starting from the nucleus and $E_1 < E_2 < E_3$ …



Bohr's Atomic Model Contd…

(b) The electron in this orbit does not radiate energy, unless a transition to a different orbit occurs. That is why these are called **stationary orbit**. When an electron jumps **from lower energy level to higher** it **absorbs** energy and *vice versa*.

$$E_2 - E_1 = \Delta E = h\nu$$

Total energy of an electron = ***kinetic energy*** + ***potential energy***

$$= \frac{1}{2}mv^2 - \frac{1}{4\pi\epsilon_0} \times \frac{ze^2}{r^2} \; ;$$

ϵ_0 = permitivity of medium = $8.85 \times 10^{-12}\ J^{-1}C^2m^{-2}$

Centrefugal force = centrepetal force

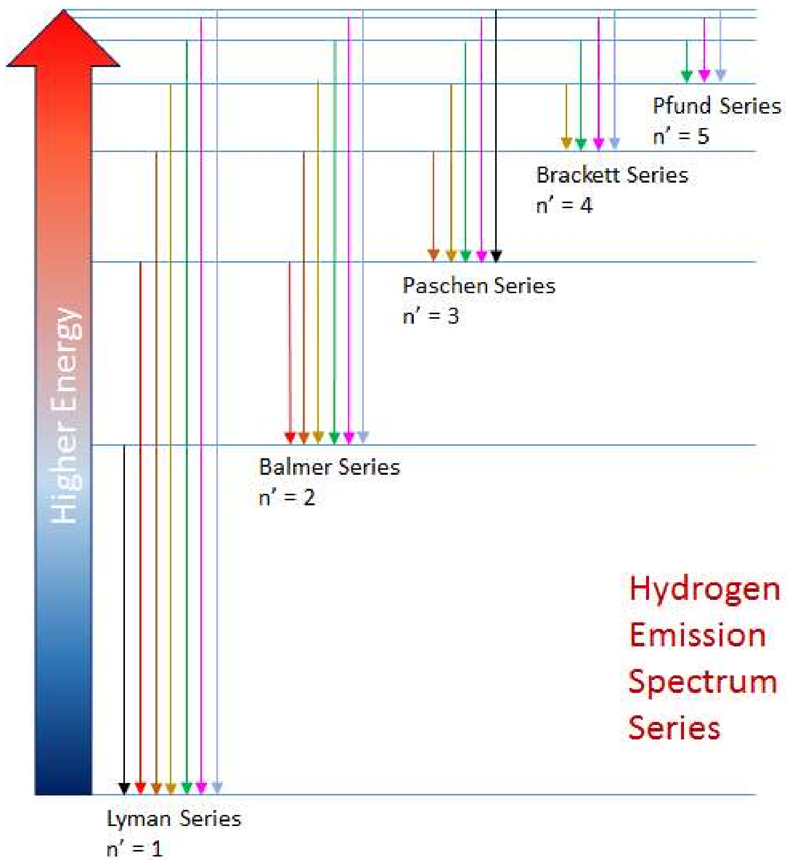
$$\Rightarrow \frac{mv^2}{r} = \frac{1}{4\pi\epsilon_0}\frac{Ze^2}{r^2} \Rightarrow v^2 = \frac{Ze^2}{4\pi\epsilon_0 rm}$$

Now, $r = \frac{nh}{2\pi mv}$

Hence,

$$E_n = -\frac{Z^2me^4}{8{\epsilon_0}^2h^2n^2}$$

$$r_n = \frac{n^2h^2}{4\pi^2me^2Z} = \frac{0 \cdot 529n^2}{Z}\text{Å} \quad \text{(Bohr radius)}$$



$$E_n = -\frac{Z^2me^4}{8\epsilon_0{}^2h^2n^2} = -2\cdot 18\times 10^{-18}\times\frac{Z^2}{n^2}J = -13\cdot 6\times\frac{Z^2}{n^2}\,\mathrm{eV}$$

Now,

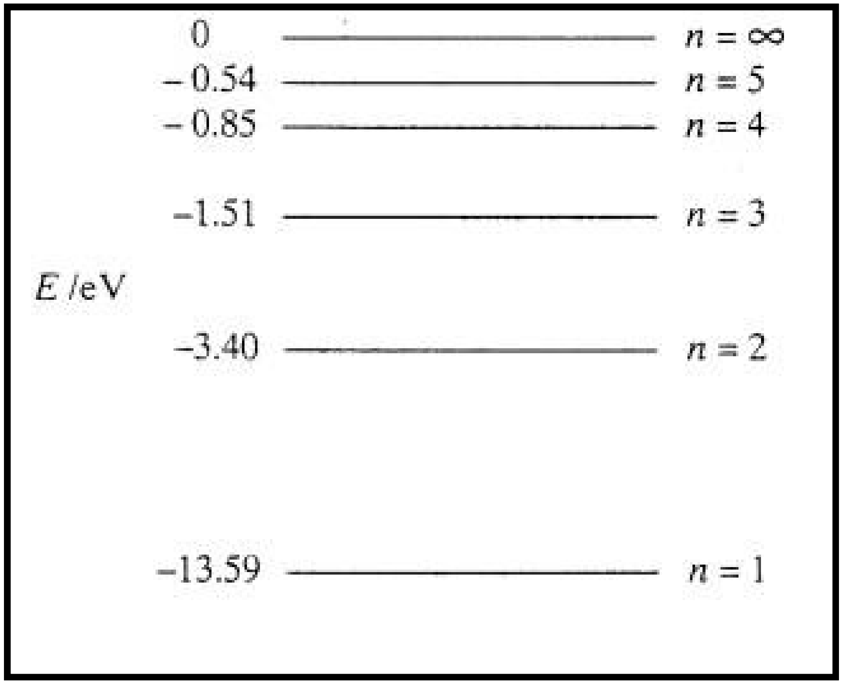
$$\Delta E = E_2 - E_1 = \frac{Z^2me^4}{8\epsilon_0^2h^2}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$\Rightarrow h\nu = \frac{Z^2me^4}{8\epsilon_0^2h^2}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$\Rightarrow \frac{hc}{\lambda} = \frac{Z^2me^4}{8\epsilon_0^2h^2}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$\Rightarrow \frac{1}{\lambda} = \frac{Z^2me^4}{8\epsilon_0^2h^3c}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$R_H = \frac{Z^2me^4}{8\epsilon_0^2h^3c} = 109{,}677\ \mathrm{cm^{-1}}$$



Energy levels in the diagram for a hydrogen atom

Achievements of Bohr's Atomic Model

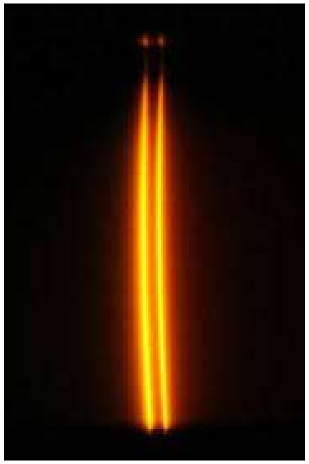
Bohr's atomic model explains the stability of an atom: According to Bohr's theory, the electron does not lose energy as long as it revolves in a particular orbit. Also, it cannot jump from the first orbit to the lower orbit as there is no orbit less than one. Thus, gradual loss of energy by the electron is not possible. Hence, the atom is stable. Thus, it removes the limitation of the Rutherford model of atoms.

Bohr's theory has explained the atomic spectra of hydrogen atoms: According to Bohr's model of atoms, electrons in an atom can have only certain definite energy levels. When the electron is present at the lowest possible energy level, it is said to be in the ground state. When the energy is supplied from some external source, the electron may absorb energy and jump to a higher energy level.

Explanation for the simultaneous appearance of many spectral lines in the hydrogen spectrum: Although the hydrogen atom has only one electron (which can be excited to one higher orbit at a time) yet hydrogen spectrum consists of several spectral lines in different series of hydrogen spectrum such as Lyman, Balmer, Paschen, Brackett and Pfund series.

Limitations of Bohr's Atomic Model

- Bohr's theory can only explain the spectra of hydrogen atoms like one electronic system e.g. He^{+}, Li^{2+}, etc. However, the theory could not explain the atomic spectra of elements having more than one electron.
- This theory cannot explain the fine-line structure of the spectrum of the hydrogen atom. Because if the hydrogen spectrum is observed with a high-resolution spectrometer, it is found that, some of the lines split into fine lines.
- It could not explain the **Zeeman effect** when the **spectral lines are split into closely spaced lines under the influence of a magnetic field** and the **Stark effect** when the **spectral lines get split into fine lines under the influence of an electric field**.
- There was no justification for the assumption that the electron can only be rotated only in those orbits in which the angular momentum of the electron should be $\frac{nh}{2\pi}$ Where h=plank's constant. **(de Broglie's hypothesis)**
- Bohr's Theory contradicts the **Heisenberg Uncertainty Principle** which states that ***"It is impossible to determine both the position and momentum (or velocity) of an electron simultaneously and accurately"***.



Sodium D lines
λ = 589 and 589.6 nm