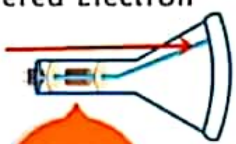
HISTORY OF ATOMIC MODEL

1885

Johann Balmer derived a formula for mathematically predicting hydrogen spectrum.

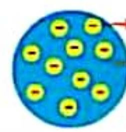
1897

J J Thomson discovered Electron



1904

J J Thomson proposed plum pudding model



1911

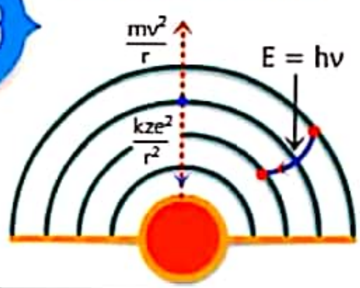
Rutherford proposed a model where positive charge is at the center, and electron moves around in a spiral path and loses energy.

1913

Bohr's Atomic Model

$$r = 0.529 \times \frac{n^2}{z} \mathring{A}$$

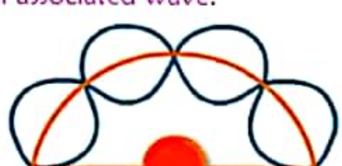
$$\frac{kze^2}{r^2} = \frac{mv^2}{r}$$



- Worked with J J Thomson and found flaws in his theory.
- He proposed electron revolves around nucleus in orbits.
- Electron is stabilized by centripetal and electrostatic forces.
- Electrons don't lose energy in an orbit.
- Electron loses or gains energy by moving across orbits.
- He proved Balmer was right by derving his formula theoretically.
- Only applicable for one electron systems.
- Failed to predict dual nature of electron.

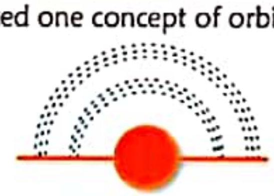
1923

De Broglie introduced the concept of dual nature in electrons. He used Einstein's $E = mc^2$ and proposed any moving particle or object has an associated wave.



1925

Erwin Schrodinger developed electron cloud model using de Broglie and Bohr's atomic model. He and Heisenberg determined the regions in which electron would be likely found. He introduced one concept of orbitals.



HYDROGEN SPECTRUM

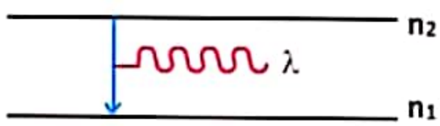
Bohr's model

Niels Bohr proposed a model for the hydrogen atom that explained the spectrum of the hydrogen atom. The Bohr model was based on the following assumptions.

- The electron in a hydrogen atom travels around the nucleus in a circular orbit.
- The energy of the electron in an orbit is proportional to its distance from the nucleus. The further the electron is from the nucleus, the more energy it has.
- Only a limited number of orbits with certain energies are allowed. In other words, the orbits are quantized.
- The only orbits that are allowed are those for which the angular momentum of the electron is an integral multiple of Planck's constant divided by 2π.

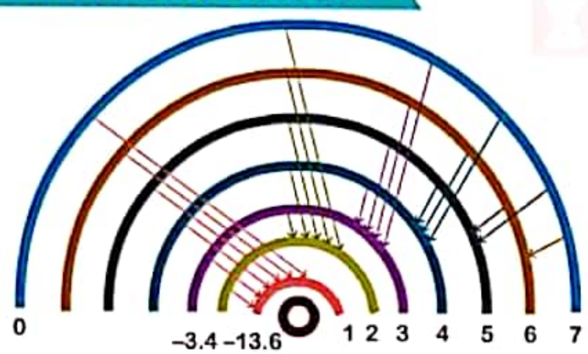
$$L = \frac{nh}{2\pi} \text{ (where h = planck's constant)}$$

- Light is absorbed when an electron jumps to a higher energy orbit and emitted when an electron falls into a lower energy orbit.
- The energy of the light emitted or absorbed is exactly equal to the difference between the energies of the orbits
- When electron in an excited atom comes back from higher energy level (n_2) to lower energy level (n_1) then it emits a photon, having energy equal to difference in energy levels.
- $h\nu = \Delta E = E_{n_2} - E_{n_1}$



Wavelength or wave no. of any line of any one electron species can be calculated as $\frac{1}{\lambda} = R_H Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$, $\frac{hc}{\lambda} = \Delta E$

Different series



Successes

- Combining successfully Rutherford's solar system's model, with the **Planck hypothesis** on the quantified energy states at atomic level and Einstein's photons
- Explaining the atomic emission and **absorption spectra**
- Explaining the general features of the **periodic table**
- First working model for the **atom**

LYMAN SERIES	BALMER SERIES	PASCHEN SERIES	BRACKETT SERIES	PFUND SERIES	HUMPHREY SERIES
$n_f = 1$	$n_f = 2$	$n_f = 3$	$n_f = 4$	$n_f = 5$	$n_f = 6$
$n_i = 2,3,4,5,...$	$n_i = 3,4,5,6,...$	$n_i = 4,5,6,7,...$	$n_i = 5,6,7,...$	$n_i = 6,7,...$	$n_i = 7,8,...$
ULTRAVIOLET	VISIBLE	INFRARED	INFRARED	FAR INFRARED	FAR INFRARED

QUANTUM NUMBER
THE ELECTRON'S ADDRESS

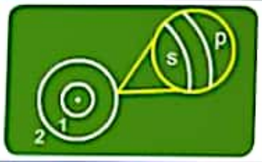
GPS is used to track anyone at any place on the earth.

n, ℓ, m Similarly Quantum Numbers are used to identify the position of an electron in an atom.

QUANTUM NUMBER

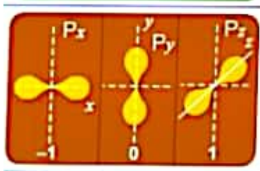
Principal Quantum Number (n)

Represents the orbit number in an atom.
It is denoted by letter 'n'.



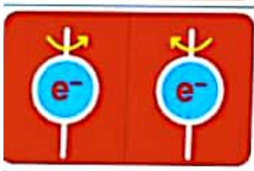
Azimuthal Quantum Number (ℓ)

Represents the shape of an orbital in atom.
it is denoted by letter 'ℓ' and its value vary from 0 to 'n'.



Magnetic Quantum Number(m_ℓ)

Represents the orientation of an orbital in the space. it is denoted by letter 'm_ℓ' and its value vary from 'ℓ' to '-ℓ'.



Spin Quantum Number(m_s)

Represents the spin of an electron.
It is denoted by m_s and each electron has an orbital either $\frac{1}{2}$ or $-\frac{1}{2}$

DIFFERENCE BETWEEN BOTH ATOMIC MODELS

Bohr's model is a 2-Dimensional model. Therefore he used only Principal quantum number (n) to identify the position of an electron in an atom.

Schrodinger's model is a 3-Dimensional model. Therefore he used n, ℓ, m to identify the position of an electron in an atom.

Hide Something at $n = 1$, $\ell = 1$, $m_\ell = 0$, $m_s = \frac{1}{2}$

Quantum numbers have some restrictions. It's not possible to find an electron at every possible combination of n, ℓ, m, s. So you will never find an electron at above point.

No two electrons in an atom have same Quantum Number.

2Px

$n = 2$	$\ell = 1$	←↑ ↓→	$n = 2$	$\ell = 1$
$m = 0$	$S = \frac{-1}{2}$		$m = 0$	$S = \frac{1}{2}$