

BIOPHYSICS LECTURE

CHAPTER 2

LESSON 1: ATOMIC MODELS

J.J Thompson's "Plum Pudding Model" of the atom failed to explain experimental results associated with the atomic structure of elements.

Rutherford's atomic model is described by the positively charged particles where most of the mass of an atom was concentrated in the nucleus. Electrons surround the nucleus of an atom and revolve around it with at high speeds in circular paths called orbits. This model did not explain how electrons were distributed in the space surrounding the nucleus.

There is a need for a better model to explain why the electrons are not attached to the protons. And this is when the Bohr's model of the hydrogen atom came along.

HISTORY OF THE ATOM

- Atomic Theory – everything is made up of tiny particles that can be break up, separated by each other in an empty space
- 1800's – John Dalton described atoms as 'solid spheres'
- 1897 – J.J Thompson Plum Pudding Model; couldn't be solid spheres
- 1909 – Ernest Rutherford; Alpha particles, thin sheet of gold; proves JJ's theory wrong; Nuclear Model; flaw is that they didn't see anything stopping the atoms
- 1913 – Niels Bohr; suggested that electrons orbited the electrons; they were held in shells; this idea was important bcs the shells prevent the collapse
- Ernest Rutherford – Protons
- James Chadwick – Neutrons

BOHR ATOMIC MODEL

- Described the atoms as positively charged nucleus comprised of protons and neutrons surrounded by negatively charged electron orbiting the nucleus in atomic shells

$$K_n = \frac{1}{2} \cdot \frac{mv^2}{4\pi\epsilon_0 r^2} = \frac{22.069eV}{n^2} \quad \text{(Kinetic energies in Bohr model)}$$

$$U_n = -\frac{1}{4\pi\epsilon_0} \cdot \frac{me^2}{4\pi\epsilon_0 r^2} = -\frac{22.069eV}{n^2} \quad \text{(Potential energies in Bohr model)}$$

The sum of these energies is:

$$E_n = -\frac{1}{4\pi\epsilon_0} \cdot \frac{me^2}{8\pi\epsilon_0 r^2} \quad \text{(Total energies in Bohr model) or}$$

Where,

- m = mass of an electron ($m = 1.019 \times 10^{-31}$ kg)
- e = charge of the electron (1.602×10^{-19} Coulombs)
- h = Planck's constant (6.626×10^{-34} J.s = 4.136×10^{-15} eV.s)
- ϵ_0 = permittivity constant (8.854×10^{-12} F/m)

In this model, the electrons in an atom can only be found in certain discrete energy states. These states are associated with the orbits or shells of electrons in an atom. Only certain photon energies are allowed when electrons jump down from higher levels to lower levels, producing the hydrogen spectrum.



When an electron moves from a higher energy level to a lower one, a photon is emitted. The energy of the emitted photon is expressed by the equation below.

$$E = hf = \frac{hc}{\lambda}$$

where,

$f = \frac{c}{\lambda}$ is the frequency in Hertz

c = 3.00×10^8 m/s is the speed of light in vacuum

λ is the wavelength of light in meter

The Bohr model worked for explaining the hydrogen atom and other single electron system such as He^+ but...

- is unable to explain spectra of atoms with more than one electron
- there is no way of explaining why some lines is more intense than the other or

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Michael Seilmaier



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