

Chapter 06

Structure of atom

Planck's Quantum Theory:

Bohr's model and its limitations, the need for a new model from the past model. Planck proposed that energy is radiated in very minute and discrete quantised amounts or packets rather than in a continuous unbroken wave. That energy is not released continuously.

Planck called the packets of energy as Quanta and he was able to determine that the energy of each quantum is equal to the frequency of the radiation multiplied by a universal constant that he derived. It is now known as Planck's constant.

Example:

Energy is released as $h\nu$, $2 h\nu$, $3 h\nu$. It goes on to $n h\nu$.

The energy for a certain frequency, E can be represented by $E = h\nu$, where h is Planck's constant.

$h = 6.626 \times 10^{-34}$ J/S and ' ν ' is the frequency of radiation

Cupric Chloride, in a watch glass and make a paste with concentrated HCl. Now take this paste on a platinum loop and heat it on a flame.

Depending on the characteristic of this copper, it emits certain color.

Example:

Salt Sodium Chloride, you heat it, you get certain lines and those lines are unique to sodium.

Barium Chloride, you get some particular lines that are unique to Barium.

So just like fingerprints, atoms also have or metals also have particular line spectra.

Bohr's model of hydrogen atom and its limitations:

Neil's Bohr proposed that electrons in an atom occupy stationary orbits or stationary states of particular energy at different distances from nucleus.

When an electron jumps from lower energy state, which is called as ground state to any higher energy states, we call the higher energy states as excited states. So basically, when an electron jumps from ground state to excited state, it absorbs energy. Similarly, when it jumps from higher energy to lower energy state, it emits some energy.

The energy absorbed or emitted is equal to the difference between the energy of two levels.

Consider, E_1 and E_2 are two different levels. Electron jumps from E_1 to E_2 .

The energy is absorbed by the electrons is equals to $E_2 - E_1$, which is equals to ' $h\nu$ ' according to the Planck's equation.

Bohr's model explains the line spectra absorbed in the case of hydrogen atom. It is successful model.

Bohr's model failed to account for a state and also for energy levels of atoms which have more than one electron.

Bohr's Sommerfeild model of an atom, to account for structure or splitting of line spectra, Somerfield model modified by Bohr's atomic model by adding elliptical orbits. First orbit is a circular orbit.

For example, an atom has three or four orbits but one comes after the first are elliptical in shape. Such that the nucleus of an atom is one of the principle of focci of these elliptical orbits.

Bohr somerfield model, though successful in accouting for the fine line structures of the hydrogen atomic spectra, it doesn't provide satisfactory picture of the structure of atom in general.

This model failed to account for the atomic spectra of atoms with more than one electron.

