Bohr's Theory of Hydrogen Atom and Hydrogen Spectra

Mariia Moiseeva

25, January 2016

1 Bohr's Theory of electron

In the Bohr model, the most stable, lowest energy level is found in the innermost orbit. This first orbital forms a shell around the nucleus and is assigned a principal quantum number (n) of n=1. Additional orbital shells are assigned values n=2, n=3, n=4, etc.

As electrons move farther away from the nucleus, they gain potential energy and become less stable. Atoms with electrons in their lowest energy orbits are in a "ground" state, and those with electrons jumped to higher energy orbits are in an "excited" state. Atoms may acquire energy that excites electrons by random thermal collisions, collisions with subatomic particles, or by absorbing a photon. Of all the photons (quantum packets of light energy) that an atom can absorb, only those having an energy equal to the energy difference between allowed electron orbits will be absorbed. Atoms give up excess internal energy by giving off photons as electrons return to lower energy (inner) orbits.

According to the Bohr model, when an electron is excited by energy it jumps from its ground state to an excited state (i.e., a higher energy orbital). The excited atom can then emit energy only in certain (quantized) amounts as its electrons jump back to lower energy orbits

2 Bohr's Hydrogen atom

Niels Bohr suggested that the problem about hydrogen spectrum can be solved if we can make some assumptions. According to classical theory, the frequency of the electromagnetic waves emitted by a revolving electron is equal to the frequency of revolution. As the electrons radiate energy, their angular velocities would change continuously and they would emit a continuous spectrum against line spectrum actually observed. So, Bohr concluded that even if electromagnetic theory successfully explained the macroscopic phenomenon, it could not be applied to explain microscopic phenomenon, that in atomic scale. So he made bold suggestions called as Bohr's postulates.

Every atom consists of nucleus and suitable number of electrons revolved around the nucleus in circular orbits.

Electrons revolved only in certain non-radiating orbits called stationery orbits for which the total angular momentum is an integral multiple of h/2p where h is plank's constant, i.e

$$L = mvr = \frac{nh}{2\pi}$$

where n=1,2,3...

The Hydrogen atom can emit different wavelengths of light depending on the initial and final energy levels of the transition. This photon of energy can be calculated with formula:

$$E = R \times (\frac{1}{n_f^2} - \frac{1}{n_i^2})$$

$$R = 8.31 \frac{J}{Kmol}$$

3 Hydrogen Spectra

Separation energy of the electron from nucleus

$$E = -2.18 \times 10^{18} Joules$$

1) Lyman series (n = 1). The series is named after its discoverer, Theodore Lyman, who discovered the spectral lines from 1906–1914. All the wavelengths in the Lyman series are in the ultraviolet band.

2) Balmer series (n = 2). Named after Johann Balmer, who discovered the Balmer formula, an empirical equation to predict the Balmer series, in 1885. Balmer lines are historically referred to as "H-alpha", "H-beta", "H-gamma" and so on, where H is the element hydrogen.Four of the Balmer lines are in the technically "visible" part of the spectrum, with wavelengths longer than 400 nm and shorter than 700 nm. Parts of the Balmer series can be seen in the solar spectrum.

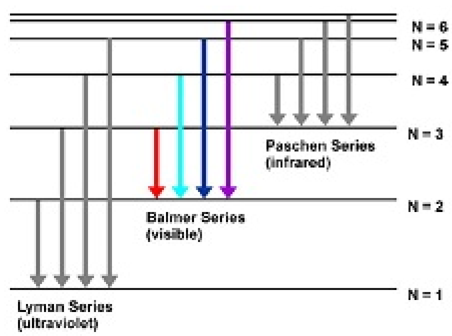

Figure 1: Balmer series

3) Paschen series (n = 3).Named after the German physicist Friedrich Paschen who first observed them in 1908. The Paschen lines all lie in the infrared band.This series overlaps with the next (Brackett) series, i.e. the shortest line in the Brackett series has a wavelength that falls among the Paschen series. All subsequent series overlap.