Bohr's Theory

Chao Yang Cheng

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1 Introduction

The Bohr model is a relatively primitive model of the hydrogen atom, compared to the valence shell atom. As a theory, it can be derived as a first-order approximation of the hydrogen atom using the broader and much more accurate quantum mechanics.
According to classical electrodynamics, a charge, which is subject to centripetal acceleration on a circular orbit, should continuously radiate electromagnetic waves. Thus, because of the loss of energy, the electron should spiral into the nucleus very soon. By contest, an electron in Bohr's model emits no energy, as long as its energy has one of the above-mentioned values. However, an electron which is not in the lowest energy level (n = 1), can make a spontaneous change to a lower state and thereby emit the energy difference in the form of a photon (particle of light). By calculating the wavelengths of the corresponding electromagnetic waves, one will get the same results as by measuring the lines of the hydrogen spectrum.

2 The theory of the observed spectra of emitted light of excited hydrogen

When a gaseous hydrogen atom in its ground state is excited by an input of energy, its electron is 'promoted' from the lowest energy level to one of higher energy. The atom does not remain excited but re-emits energy as electromagnetic radiation. This is as a result of an electron 'falling' from a higher energy level to one of lower energy. This electron transition results in the release of a photon from the atom of an amount of energy (E = hn) equal to the difference in energy of the electronic energy levels involved in the transition. In a sample of gaseous hydrogen where there are many trillions of atoms all of the possible electron transitions from higher to lower energy levels will take place many times. A prism can now be used to separate the emitted electromagnetic radiation into its component frequencies (wavelengths or energies). These are then represented as spectral lines along an increasing frequency scale to form an atomic emission spectrum.

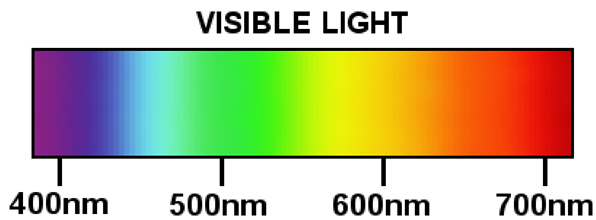

Figure 1: Visible light

3 The 1st 3 hydrogen spectra

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.

n	λ, vacuum (nm)
2	121.57
3	102.57
4	97.254
5	94.974
6	93.780
∞	91.175

Figure 2: Lyman Series

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. H-alpha is an important line used in astronomy to detect the presence of hydrogen.

n	λ, air (nm)
3	656.3
4	486.1
5	434.0
6	410.2
7	397.0
∞	364.6

Figure 3: 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.

n	λ, air (nm)
4	1875
5	1282
6	1094
7	1005
8	954.6
9	922.9
∞	820.4

Figure 4: Paschen Series

4 Relevant constant and formulae

Constant:
$c = 3 \times 10^8 \frac{m}{s}$
$M_e = 9.1 \times 10^{-31} kg$
$e = 1.6 \times 10^{-19} C$
$k_e = 9.0 \times 10^9 \frac{Nm^2}{C^2}$
$h = 6.63 \times 10^{-34} J$
$h_{bar} = \frac{h}{2\pi}$

Formulae:
$F_E = \frac{e^2 k_e}{r^2}$
$KE = \frac{1}{2} m_e v^2$
$PE = \frac{k_e q_1 q_2}{r} = \frac{K_E e^2}{r}$
$L = rmv = nh_{bar}$

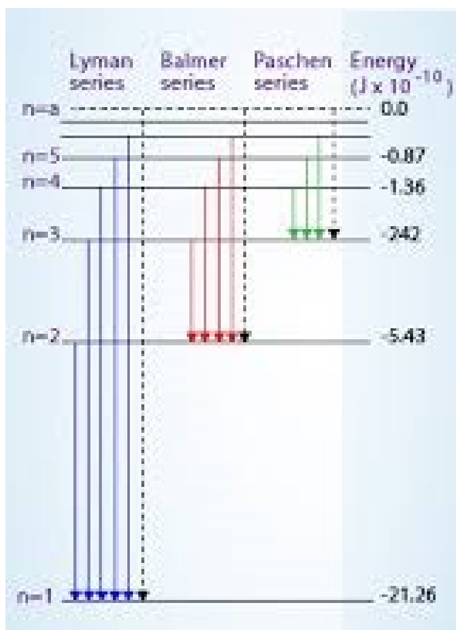

Figure 5: The 1st 3 Spectra