Bohr's Theory of Hydrogen Atom and Hydrogen Spectra

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1 Introduction

In 1913, Danish physicist Niels Bohr introduced the atomic model of Hydrogen, which he described as a positively charged nucleus consisting of protons and neutrons, surrounded by a negatively charged electron cloud. The atom is held together by electrostatic forces between the positive nucleus and negative surroundings. The negatively charged electrons revolve about the positively charged atomic nucleus because of the attractive electrostatic force according to Coulomb's law.

2 Hydrogen Energy Levels

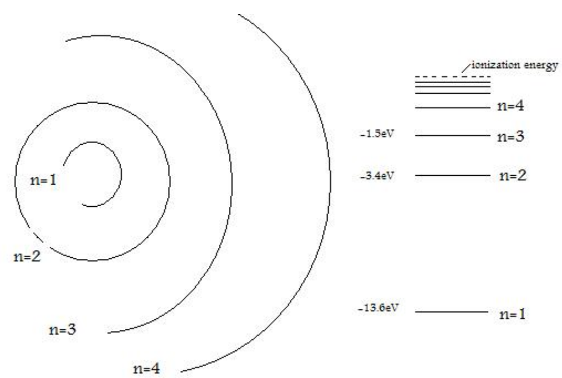

Figure 1: The shell structure of a Hydrogen atom, where each shell is associated with principle quantum number n

3 The Electromagnetic Spectrum

The electromagnetic (EM) spectrum is the range of all types of EM radiation. Radiation is energy that travels and spreads out as it goes. The EM spectrum comprises of visible light, radio waves microwaves, infrared light, ultraviolet light, X-rays and gamma-rays.

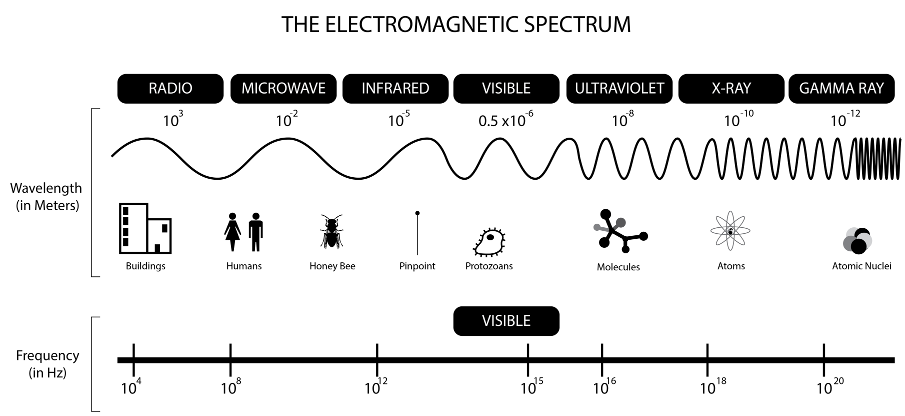

Figure 2: Different types of radiation in the electromagnetic spectrum

4 Excitation

Excitation is the addition of a discrete amount of energy (called excitation energy) to a system, such as an atomic nucleus, an atom, or a molecule that results in its alteration, ordinarily from the condition of lowest energy (ground state) to one of higher energy (excited state).

5 The Rydberg Formula

The Rydberg formula explains the different energies of transition that occur between energy levels. When an electron moves from a higher energy level to a lower one, a photon is emitted. The Hydrogen atom can emit different wavelengths of light depending on the initial and final energy levels of the transition. It emits a photon with energy equal to the difference of square of the final (n_f) and initial (n_i)energy levels.

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

Planck's constant => h= 6.63 $\times 10^{-34}$m^{2}kg/s

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

The combination of these two equations produces the Rydberg Formula.

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

The Rydberg Constant (R) = 1.097 $\times 10^7$m^{-1}

6 Hydrogen Spectrum

The movement of electrons between these energy levels produces a spectrum. The Balmer equation is used to describe the four different wavelengths of Hydrogen which are present in the visible light spectrum: 656, 486, 434, and 410nm. These correspond to the emission of photons as an electron in an excited state transitions down to energy level n=2.

Example: In order to get wavelength from n=3 to n=2, we can use the Rydberg formula.

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

$\frac{1}{\lambda} = (1.097 \times 10^7 1/\mathrm{m})(\frac{1}{2^2} - \frac{1}{3^2})$

$\frac{1}{\lambda} = 1523611$

$\lambda = 656 \times 10^{-9}\mathrm{m}$

$\lambda = 656\mathrm{nm}$

Wavelength (nm)	Relative Intensity	Transition	Color or region of EM spectrum
Lymann Series			
93.782	...	6 -> 1	UV
94.976	...	5 -> 1	UV
97.254	...	4 -> 1	UV
102.583	...	3 -> 1	UV
121.566	...	2 -> 1	UV
Balmer Series			
383.5384	5	9 -> 2	Violet
388.9049	6	8 -> 2	Violet
397.0072	8	7 -> 2	Violet
410.174	15	6 -> 2	Violet
434.047	30	5 -> 2	Violet
486.133	80	4 -> 2	Bluegreen (cyan)
656.272	120	3 -> 2	Red
656.2852	180	3 -> 2	Red
Paschen Series			
954.62	...	8 -> 3	IR
1004.98	...	7 -> 3	IR
1093.8	...	6 -> 3	IR
1281.81	...	5 -> 3	IR
1875.01	...	4 -> 3	IR

Figure 3: Measured Hydrogen Spectrum of Lyman, Balmer, and Pachen series

References

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