

Bohr's theory of hydrogen atom

Nghi Nguyen

January 2016

1 Introduction

In this research paper, I'm going to focus on examining and analyzing the definition as well as the structure of Bohr's hydrogen atom. Hence, I can certainly link it from

2 What is Bohr's hydrogen atom and when was it introduced?

Neils Bohr introduced the atomic hydrogen model in 1913. It's described as a positively charged nucleus, comprised of protons and neutrons, surrounded by a negatively charged electron cloud. In the model, electrons orbit the nucleus in atomic shells. The atom is held together by electrostatic forces between the positive nucleus and negative surrounding.

3 Bohr's theory of hydrogen atom under perception of Physics

Quantized Energy States: The electrons in free atoms can be found in only certain discrete energy states. These sharp energy states are associated with the orbits or shells of electrons in an atom (which is a hydrogen atom). The Bohr Model successfully predicted the energies for the hydrogen atom, but had significant failures that were connected by solving the Schrodinger equation for the hydrogen atom.

Angular Momentum Quantization: In the Bohr model, the wavelength associated with the electron is given by the DeBroglie relationship:

$$\lambda = \frac{h}{mv}$$

The standing wave condition that circumference= whole number of wavelength. In the hydrogenic case, the number n is the principal quantum number.

$$2\pi r = n\lambda_n$$

Those can be combined to get an expression for the angular momentum of the electron in orbit.

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

Bohr Orbit: Combining the energy of the classical electron orbit with the quantization of angular momentum, the Bohr approach yields expressions for the electrons orbit radio and energies.

$$\frac{mv^2}{2} = \frac{(mvr)^2}{2mr^2} = \frac{n^2h^2}{8\pi^2}$$

4 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. These wavelengths are at 656, 486, 434, and 410nm. These correspond to the emission of photons as an electron in an excited state transitions down to energy level $n=2$. The Rydberg formula, below, generalizes the Balmer series for all energy level transitions.

5 Triumphs and Limitations of the Bohr model

Bohr did what no one had been able to do before. Not only did he explain the spectrum of hydrogen, he correctly calculated the size of the atom from basic physics. Some of his ideas are broadly applicable. Electron orbital energies are quantized in all atoms and molecules. Angular momentum is quantized. The electrons do not spiral into the nucleus, as expected classically (accelerated charges radiate, so that the electron orbits classically would decay quickly, and the electrons would sit on the nucleus—matter would collapse). These are major triumphs.

The Bohr Model was an important step in the development of atomic theory. However, it has several limitations.

It is in violation of the Heisenberg Uncertainty Principle.

The Bohr Model considers electrons to have both a known radius and orbit, which is impossible according to Heisenberg.

The Bohr Model is very limited in terms of size. Poor spectral predictions are obtained when larger atoms are in question.

It cannot predict the relative intensities of spectral lines.

It does not explain the Zeeman Effect, when the spectral line is split into several components in the presence of a magnetic field.

The Bohr Model does not account for the fact that accelerating electrons do not emit electromagnetic radiation.