The Bohr

Yunpeng Huang

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1 Introduction:Bohr model

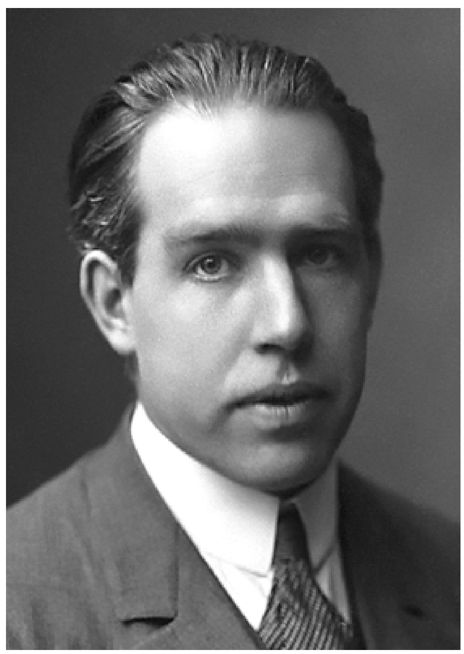

Figure 1: Bohr

In atomic physics, the RutherfordBohr model or Bohr model, introduced by Niels Bohr in 1913, depicts the atom as a small, positively charged nucleus surrounded by electrons that travel in circular orbits around the nucleussimilar in structure to the solar system, but with attraction provided by electrostatic forces rather than gravity. After the cubic model (1902), the plum-pudding model (1904), the Saturnian model (1904), and the Rutherford model (1911) came the RutherfordBohr model or just Bohr model for short (1913). The improvement to the Rutherford model is mostly a quantum physical interpretation of it. The Bohr model has been superseded, but the quantum theory remains sound.

2 Introduction:Bohr Theory of Hydrogen Atom

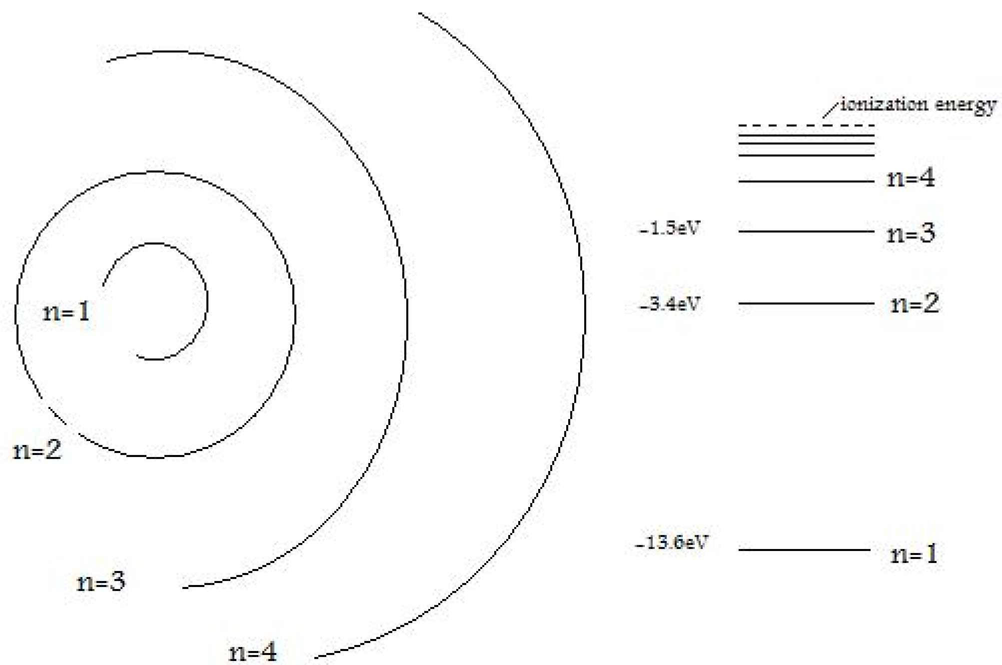

Figure 2: Atom and Energy level

Niels Bohr introduced the atomic Hydrogen model in 1913. He described it 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 surroundings.

3 Introdiction:Hydrogen Energy Levels

The Bohr model is used to describe the structure of hydrogen energy levels. The image below represents shell structure, where each shell is associated with principle quantum number n. The energy levels presented correspond with each shell. The amount of energy in each level is reported in eV, and the maxiumum energy is the ionization energy of 13.598eV.

4 Introduction:Hydrogen spectra

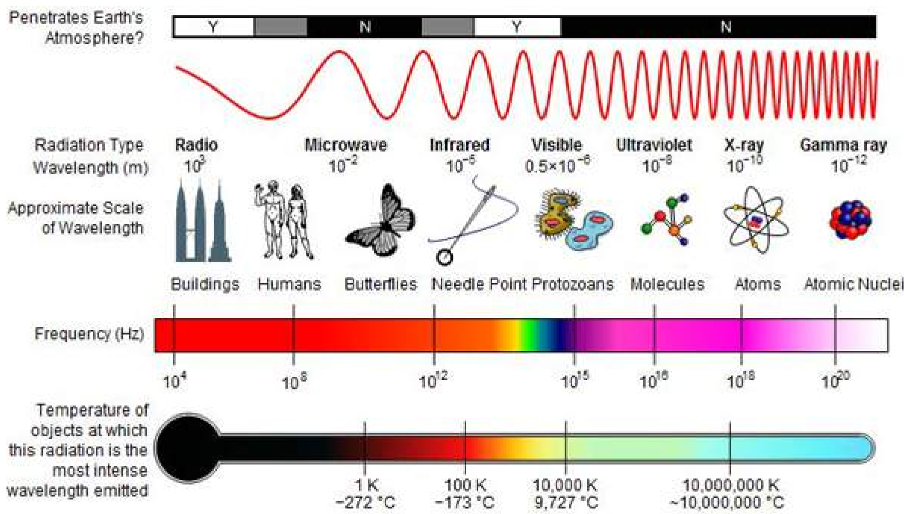

Figure 3: spectra

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. To get the Balmer lines, the Rydberg formula is used with an nf of 2.

5 Excitation

Atoms can make transitions between the orbits allowed by quantum mechanics by absorbing or emitting exactly the energy difference between the orbits. The following figure shows an atomic excitation cause by absorption of a photon and an atomic de-excitation caused by emission of a photon.

6 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 (nfnf) and initial (nini) energy levels.

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

The energy of a photon is equal to Plancks constant, h=6.626*10-34m2kg/s, times the speed of light in a vacuum, divided by the wavelength of emission.

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

Combining these two equations produces the Rydberg Formula.

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

The Rydberg Constant (R) = 10,973,731.6m110,973,731.6m1 or 1.097107m11.097107m1.