Bohr's Theory

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1 Theory

In atomic physics,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 nucleus—similar in structure to the solar system, but with attraction provided by electrostatic forces rather than gravity. The Bohr model has been superseded, but the quantum theory remains sound.

The model's key success lay in explaining the Rydberg formula for the spectral emission lines of atomic hydrogen. While the Rydberg formula had been known experimentally, it did not gain a theoretical underpinning until the Bohr model was introduced. Not only did the Bohr model explain the reason for the structure of the Rydberg formula, it also provided a justification for its empirical results in terms of fundamental physical constants.

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 and thus may be considered to be an obsolete scientific theory. However, because of its simplicity, and its correct results for selected systems, the Bohr model is still commonly taught to introduce students to quantum mechanics or energy level diagrams before moving on to the more accurate, but more complex, valence shell atom. A related model was originally proposed by Arthur Erich Haas in 1910, but was rejected.

The Bohr Model is probably familiar as the "planetary model" of the atom illustrated in the adjacent figure that, for example, is used as a symbol for atomic energy (a bit of a misnomer, since the energy in "atomic energy" is actually the energy of the nucleus, rather than the entire atom). In the Bohr Model the neutrons and protons occupy a dense central region called the nucleus, and the electrons orbit the nucleus much like planets orbiting the Sun (but the orbits are not confined to a plane as is approximately true in the Solar System). The adjacent image is not to scale since in the realistic case the radius of the nucleus is about 100,000 times smaller than the radius of the entire atom, and as far as we can tell electrons are point particles without a physical extent.

This similarity between a planetary model and the Bohr Model of the atom ultimately arises because the attractive gravitational force in a solar system and the attractive Coulomb (electrical) force between the positively charged nucleus and the negatively charged electrons in an atom are mathematically of the same form. (The form is the same, but the intrinsic strength of the Coulomb interaction is much larger than that of the gravitational interaction; in addition, there are positive and negative electrical charges so the Coulomb interaction can be either attractive or repulsive, but gravitation is always attractive in our present Universe.)

The wavelength of the emitted or absorbed light is exactly such that the photon carries the energy difference between the two orbits. This energy may be calculated by dividing the product of the Planck constant and the speed of light hc by the wavelength of the light). Thus, an atom can absorb or emit only certain discrete wavelengths (or equivalently, frequencies or energies). The frequency and the wavelength of light, emitted by the excited hydrogen, solely depends on the speed of the particle, which interacts with the system.

The emission spectrum of atomic hydrogen is divided into a number of the spectral series, with wavelengths given by the Rydberg formula. These observed spectral lines are due to the electron making transitions between two energy levels in the atom.A hydrogen atom consists of an electron orbiting its nucleus. The electromagnetic force between the electron and the nuclear proton leads to a set of quantum states for the electron, each with its own energy. These states were visualized by the Bohr model of the hydrogen atom as being distinct orbits around the nucleus.

Spectral emission occurs when an electron transitions, or jumps, from a higher energy state to a lower energy state. Because the energy of each state is fixed, the energy difference between them is fixed, and the transition will always produce a photon with the same energy.

2 Formulas

$$\Delta E = E_2 - E_1 = hv$$

$$v = \lambda f$$

$$\frac{1}{\lambda} = R(\frac{1}{n'^2} - \frac{1}{n^2})$$

3 Diagrams

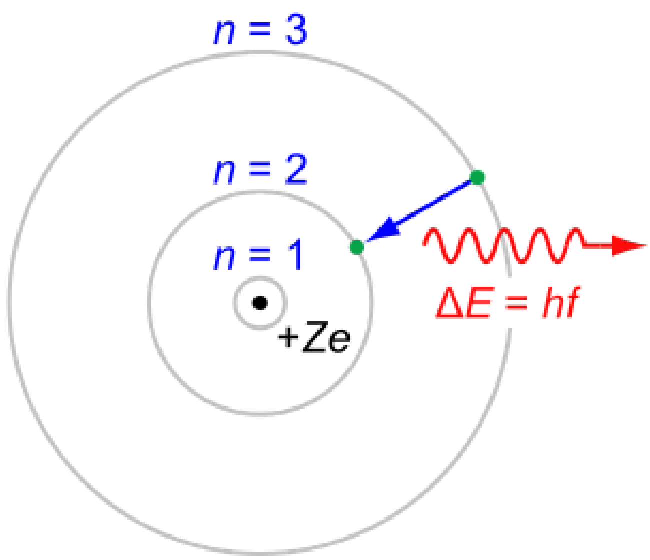
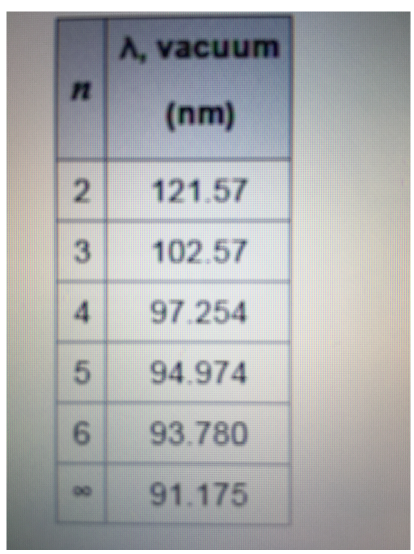


Figure 1: Bohr's model



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

Figure 2: Hydrogen Spectra

References

https://en.wikipedia.org/wiki/Bohr$_{m}odel$
http://csep10.phys.utk.edu/astr162/lect/light/bohr.html
https://en.wikipedia.org/wiki/Hydrogen$_{s}pectral_{s}eries$