2.8: Summary of Bohr's Contribution

Bohr's proposal explained the hydrogen atom spectrum, the origin of the Rydberg formula, and the value of the Rydberg constant. Specifically it demonstrated that the integers in the Rydberg formula are a manifestation of quantization. The energy, the angular momentum, and the radius of the orbiting electron all are quantized. This quantization also parallels the concept of stable orbits in the Bohr model. Only certain values of E, M, and r are possible, and therefore the electron cannot collapse onto the nucleus by continuously radiating energy because it can only have certain energies, and it cannot be in certain regions of space. The electron can only jump from one orbit (quantum state) to another. The quantization means that the orbits are stable, and the electron cannot spiral into the nucleus in spite of the attractive Coulomb force.

How might one have been so clever as to propose that angular momentum is quantized in units of $\hbar$? Possibly, by using unit analysis.

Example 2.8.1

Show that the units of Planck's constant (J s) are the same as angular momentum (mvr = kgm^2/s).

Example 2.8.2

Why do you suppose Bohr did not include the possibility of no angular momentum for the electron, i.e. n = 0?

The factor of $\frac{1}{2\pi}$ is needed to obtain the experimental value for RH from the theory. Without this factor Bohr would have calculated a value for RH that was smaller than the experimental value by a factor of $(2\pi)^2$.

Example 2.8.3

Calculate a value for R_H using fundamental Constants. Repeat the calculation assuming that angular momentum is quantized in units of h rather than $\hbar$. Show that the value you calculate differs from the value obtained by Rydberg from the hydrogen atom data by a factor of $(2\pi)^2$.

Although Bohr's ideas successfully explained the hydrogen spectrum, they failed when applied to the spectra of other atoms. In addition a profound question remained. Why is angular momentum quantized in units of $\hbar$? As we shall see, de Broglie had an answer to this question, and this answer led Schrödinger to a general postulate that produces the quantization of angular momentum as a consequence. This quantization is not quite as simple as proposed by Bohr, and we will see that it is not possible to determine the distance of the electron from the nucleus as precisely as Bohr thought. In fact, since the position of the electron in the hydrogen atom is not at all as well defined as a classical orbit (such as the moon orbiting the earth) it is called an orbital. An electron orbital represents or describes the position of the electron around the nucleus in terms of a mathematical function called a wavefunction that yields the probability of positions of the electron.

> *Bohr's ideas successfully explained the hydrogen spectrum, but they failed when applied to the spectra of other atoms.*

Bohr's idea that the absorption or emission of light corresponds to an electron jumping from one orbit to another also is not entirely accurate. When light is absorbed or emitted by an atom or molecule, the atom or molecule makes a transition from one energy state to another. If there is more than one state associated with each energy level, the energy level is said to be degenerate. Bohr's analysis produced the correct energy level spacing for the hydrogen atom and therefore could explain the luminescence spectrum, but the analysis did not identify all the possible electronic states of the hydrogen atom, as we shall see later in Chapter 8 with the Zeeman effect. The Bohr model also did not predict the electronic states of multielectron atoms. This spurred a long period of development culminating in quantum mechanics as we are studying in this text.

Contributors and Attributions

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