Bohr Paper

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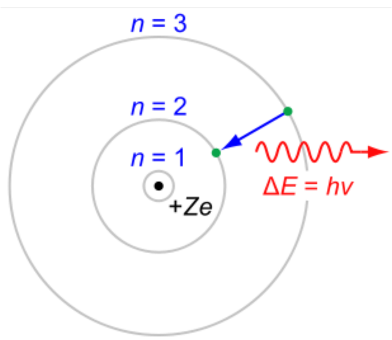
1. Introduction:

In atomic physics, the Bohr model depicts an atom as a small, positively charged nucleus surrounded by electrons. These electrons travel in circular orbits around the nucleus—similar in structure to the solar system, except electrostatic forces rather than gravity provide attraction.



The Bohr model was an improvement on the earlier cubic model (1902), the plum-pudding model (1904), the Saturnian model (1904), and the Rutherford model (1911). Since the Bohr model is a quantum-physics-based modification of the Rutherford model, many sources combine the two: the Rutherford–Bohr model.

Niels Bohr was a Danish physicist who made a fundamental contribution to our understanding of atomic structure and quantum mechanics. He made the first successful attempt at modeling the hydrogen atom by assuming that the electron executes orbital motion about the proton (i.e., the nucleus). Assuming that the only force between the electron and the proton is the electrostatic force, and applying Newton's 2nd law, he arrived at the following equation:

$$\sum F = ma \qquad \rightarrow \qquad \frac{e^2}{4\pi\epsilon_o r^2} = m\frac{v^2}{r}$$

2. Properties of Electrons:

In 1913, Bohr suggested that electrons could only have certain classical motions:

1. The electrons orbiting the nucleus.
2. The electrons can only orbit stable, no radiation, in certain orbits at a certain discrete distances from the nucleus. These orbits are related to certain energy and are also known as energy or energy

shell. In orbit, the electron acceleration and radiation do not result in loss of energy required by the classical electromagnetic theory.

3. The electrons can only gain or lose energy by jumping from a permitted trajectory others, absorb or emit electromagnetic radiation with frequencies (ν) determined by the energy difference of the levels follow Planck relationship.

Bohr's model is significant because the laws of classical mechanics apply to the motion of the electron about the nucleus only when restricted by a quantum rule. Although Rule 3 is not completely well defined for small orbits, Bohr could determine the energy spacing between levels using Rule 3 and come to an exactly correct quantum rule—the angular momentum *L* is restricted to be an integer multiple of a fixed unit:

$$L = n\,\frac{h}{2\pi} = n\hbar$$

Where n = 1, 2, 3, ... is called the principal quantum number and ħ = h/2π. The lowest value of n is 1; this gives a smallest possible orbital radius of 0.0529 nm, known as the Bohr radius. Once an electron is in this lowest orbit, it can get no closer to the proton. Starting from the angular momentum quantum rule, Bohr was able to calculate the energies of the allowed orbits of the hydrogen atom and other hydrogen-like atoms and ions.