# The Bohr Model of the Hydrogen Atom

**Unit:** Electronic Structure

**MA Curriculum Frameworks (2016):** HS-PS1-1

**Mastery Objective(s):** (Students will be able to…)

- Describe developments that led to the Bohr model of the atom.
- Describe & explain the Bohr model of the atom.
- Explain how the quantum mechanical model of the atom grew out of the Bohr model.

**Success Criteria:**

- Descriptions successfully communicate developments prior to the Bohr model that were incorporated into the model.
- Descriptions successfully communicate accurate information about the Bohr model and how it describes the behavior of atoms.

**Tier 2 Vocabulary:** model

**Language Objectives:**

- Explain scientific information about the Bohr mechanical model of the atom.

**Notes:**

## Significant Developments Prior to 1913

### Atomic Theory

Significant developments in atomic theory are described in the "History of Atomic Theory" section, which begins on page 157. The most significant advances were the discovery of the electron and the planetary model of the atom.

### Early Quantum Theory

"Old" Quantum Theory (ca. 1900): sub-atomic particles obey the laws of classical mechanics, but that only certain "allowed" states are possible.

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## Spectroscopy

Balmer Formula (1885): Swiss mathematician and physicist Johann Balmer devised an empirical equation to relate the emission lines in the visible spectrum for the hydrogen atom.

Rydberg Formula (1888): Swedish physicist Johannes Rydberg developed a generalized formula that could describe the wave numbers of all of the spectral lines in hydrogen (and similar elements).

There are several series of spectral lines for hydrogen, each of which converge at different wavelengths. Rydberg described the Balmer series in terms of a pair of integers ($n_1$ and $n_2$, where $n_1 < n_2$), and devised a single formula with a single constant (now called the Rydberg constant) that relates them.

$$\frac{1}{\lambda_{vac}} = R_H \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

The value of Rydberg's constant is $\dfrac{m_e e^4}{8\varepsilon_o^2 h^3 c} = 10\ 973\,731.6\ \text{m}^{-1} \approx 1.1 \times 10^7 \text{m}^{-1}$

where $m_e$ is the rest mass of the electron, $e$ is the elementary charge, $\varepsilon_o$ is the electrical permittivity of free space, $h$ is Planck's constant, and $c$ is the speed of light in a vacuum.

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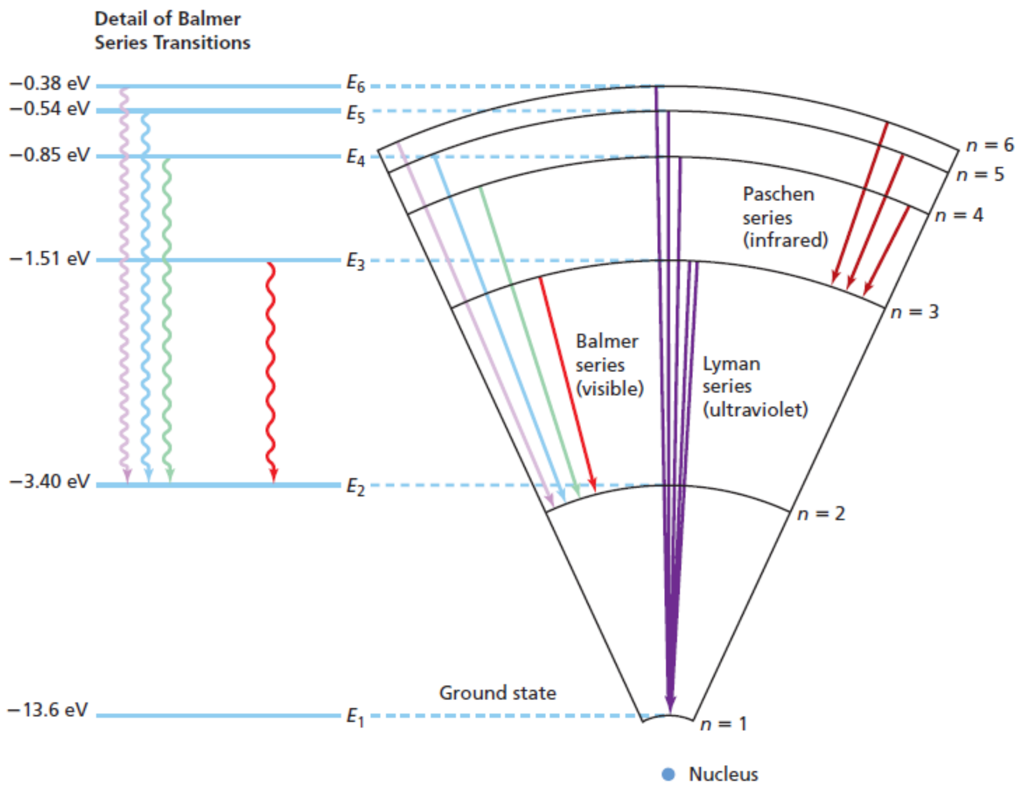
Rydberg's equation was later found to be consistent with other series discovered later, including the Lyman series (in the ultraviolet region; first discovered in 1906) and the Paschen series (in the infrared region; first discovered in 1908).

Those series and their converging wavelengths are:

| Series | Wavelength | $n_1$ | $n_2$ |
|---|---|---|---|
| Lyman | 91 nm | 1 | $2 \rightarrow \infty$ |
| Balmer | 365 nm | 2 | $3 \rightarrow \infty$ |
| Pasch7en | 820 nm | 3 | $4 \rightarrow \infty$ |

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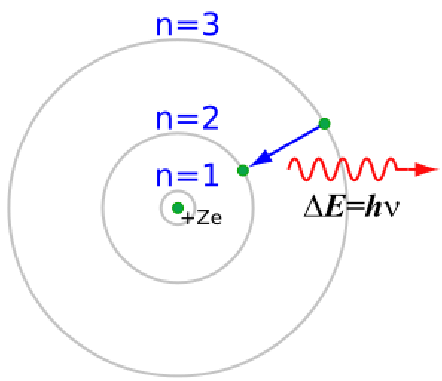
# Bohr's Model of the Atom (1913)

In 1913, Danish physicist Niels Bohr combined atomic, spectroscopy, and quantum theories into a single theory. Bohr hypothesized that electrons moved around the nucleus as in Rutherford's model, but that these electrons had only certain allowed quantum values of energy, which could be described by a quantum number (*n*). The value of that quantum number was the same *n* as in Rydberg's equation, and that using quantum numbers in Rydberg's equation could predict the wavelengths of light emitted when the electrons gained or lost energy by moved from one quantum level to another.




Bohr's model gained wide acceptance, because it related several prominent theories of the time. The theory worked well for hydrogen, giving a theoretical basis for Rydberg's equation. Bohr defined the energy released when an electron descended to an energy level using an integer quantum number (*n*) and Rydberg's constant:

$$E_n = -\frac{R_H}{n^2}$$

Bohr received the Nobel Prize in physics in 1922 for his contributions to quantum and atomic theory.

Although the Bohr model worked well for hydrogen, the equations could not be solved exactly for atoms with more than one electron, because of the additional effects that electrons exert on each other (*e.g.,* via the Coulomb force, $F_e = \frac{kq_1q_2}{r^2}$).

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