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Bohr Model of the Hydrogen Atom

Unit: Quantum and Particle Physics

NGSS Standards/MA Curriculum Frameworks (2016): N/A

AP® Physics 2 Learning Objectives/Essential Knowledge (2024): 5.B.8.1

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

- Explain how Bohr's model unified recent developments in the fields of spectroscopy, atomic theory and early quantum theory.
- Calculate the frequency/wavelength of light emitted using the Rydberg equation.
- Calculate the energy associated with a quantum number using Bohr's equation.

Success Criteria:

- Descriptions & explanations account for observed behavior.
- Variables are correctly identified and substituted correctly into the correct equation.
- Algebra is correct with correct rounding and reasonable units.

Language Objectives:

- Explain why the Bohr Model was such a big deal.

Tier 2 Vocabulary: model, quantum

Notes:

Significant Developments Prior to 1913

Discovery of the Electron (1897): English physicist J.J. Thomson determined that cathode rays were actually particles emitted from atoms that the cathode was made of. These particles had an electrical charge, so they were named "electrons" (though Thomson called them "corpuscles").

Planetary Model of the Atom (1903): Japanese physicist Hantaro Nagaoka first proposed a model of the atom in which a small nucleus was surrounded by a ring of electrons.

Discovery of the Atomic Nucleus (1909): English physicist Ernest Rutherford's famous "gold foil experiment" determined that atoms contained a dense, positively-charged nucleus that comprised most of the atom's mass.

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Spectroscopy

Balmer Formula (1885): Swiss mathematician 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 $\frac{m_e e^4}{8\varepsilon_o^2 h^3 c} = 10\,973\,731.6\text{ m}^{-1} \approx 1.1\times10^7\text{ m}^{-1}$

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

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	$3 \rightarrow \infty$
Paschen	820 nm	3	$4 \rightarrow \infty$

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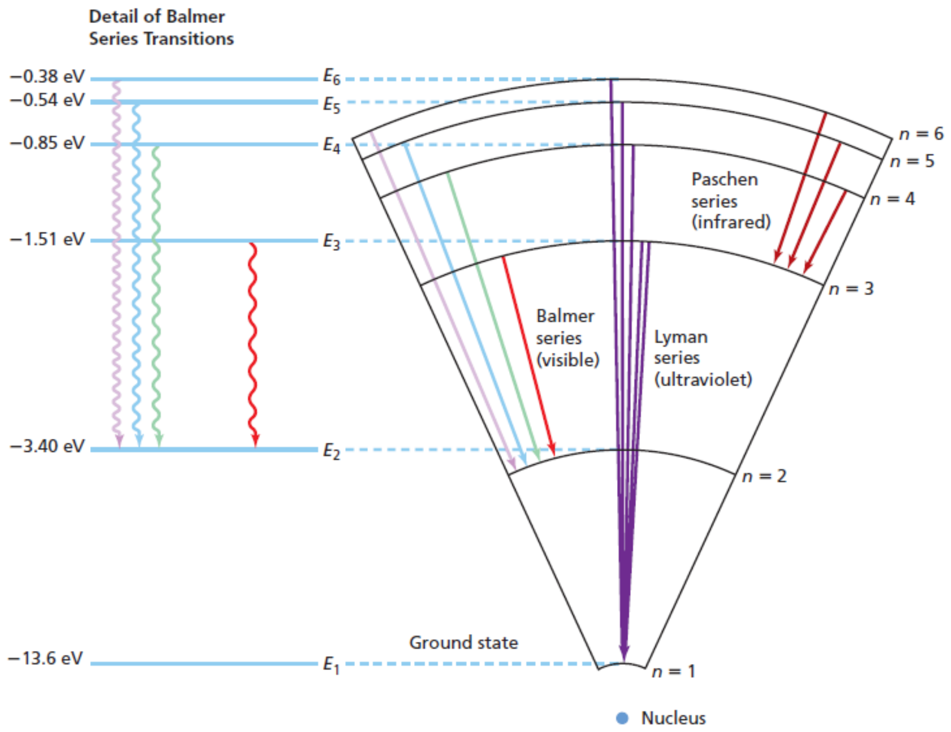
The following diagram shows Lyman, Balmer and Paschen series transitions for the hydrogen atom, from higher energy levels (n = 2 through n = 6) back to lower ones (n = 1 through n = 3):



Early Quantum Theory

quantum: an elementary unit of energy.

In 1900, German physicist Max Planck published the Planck postulate, stating that electromagnetic energy could be emitted only in quantized form, *i.e.,* only certain "allowed" energy states are possible.

Plack determined the constant that bears his name as the relationship between the frequency of one quantum unit of electromagnetic wave and its energy. This relationship is the equation:

$$E = hf$$

where:

E = energy (J)

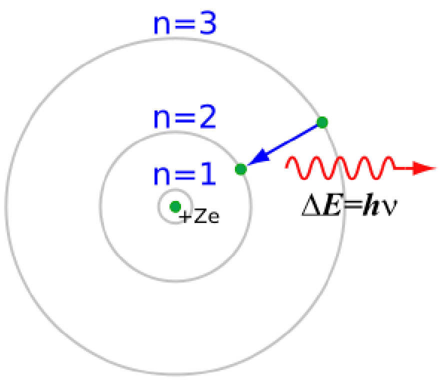
h = Planck's constant = $6.626\times10^{-34}\,\text{J}\cdot\text{s}$

f = frequency* (Hz ≡ s^{-1})

* Most physics texts use the Greek letter *v* (nu) as the variable for frequency. However, high school texts and the College Board use *f*, presumably to avoid confusion with the letter "v".

Bohr's Model of the Atom (1913)

In 1913, Danish physicist Niels Bohr combined atomic, quantum and spectroscopy theories into a single unified 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. He received a Nobel Prize in physics in 1922 for his work.

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The theory worked well for hydrogen, giving a theoretical basis for Rydberg's equation. Bohr defined the energy associated with a quantum number (n) in terms of Rydberg's constant:

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

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.,* the Coulomb force, $F_e = \frac{1}{4\pi\varepsilon_o}\frac{q_1 q_2}{r^2}$).