The Bohr Model of the Hydrogen Atom

Unit: Electronic Structure

NGSS Standards/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 140. The most significant advances were the discovery of the electron and the planetary model of the atom.

Early Quantum Theory

<u>"Old" Quantum Theory (ca. 1900)</u>: 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 $\frac{m_e e^4}{8\varepsilon_o^2 h^3 c} = 10\ 973\ 731.6\ \text{m}^{-1} \approx 1.1 \times 10^7 \text{m}^{-1}$ where

m_e is the rest mass of the electron, e is the elementary charge, ε_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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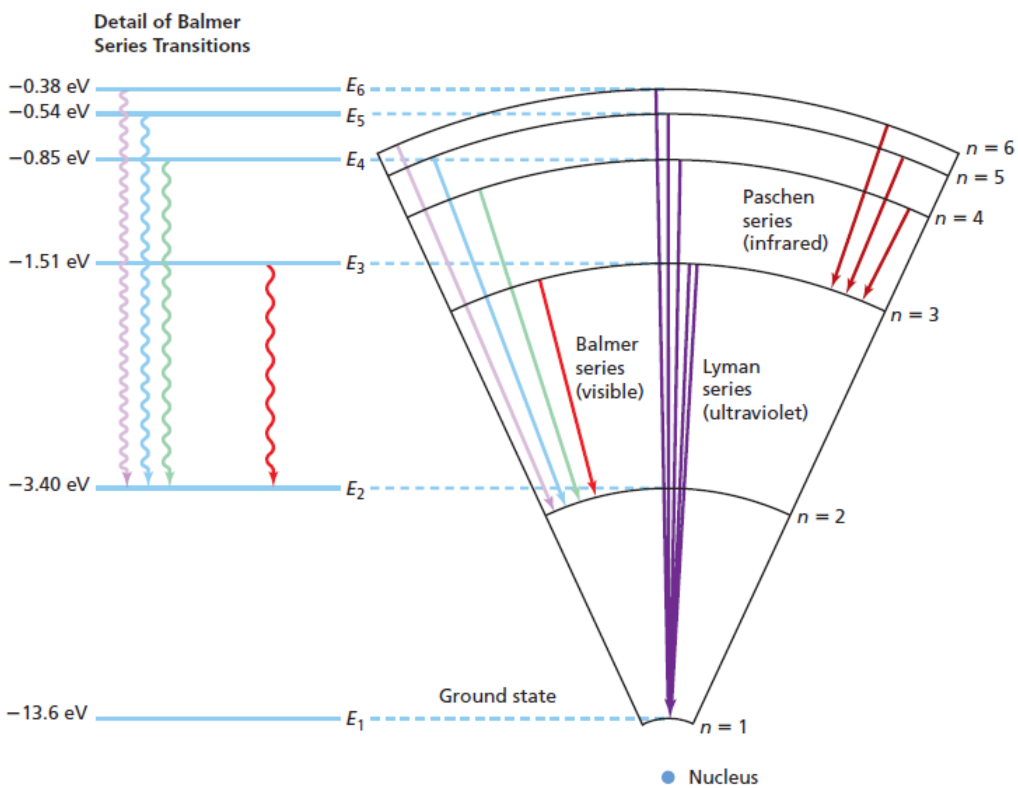
Rydberg's equation was later found to be consistent with other series discovered later, including the Lyman series (in the ultraviolet region; first discovered in 1906) and the Paschen series (in the infrared region; first discovered in 1908).

Those series and their converging wavelengths are:

Series	**Wavelength**	n_1	n_2
Lyman	91 nm	1	$2 \rightarrow \infty$
Balmer	365 nm	2	$3 \rightarrow \infty$
Paschen	820 nm	3	$4 \rightarrow \infty$



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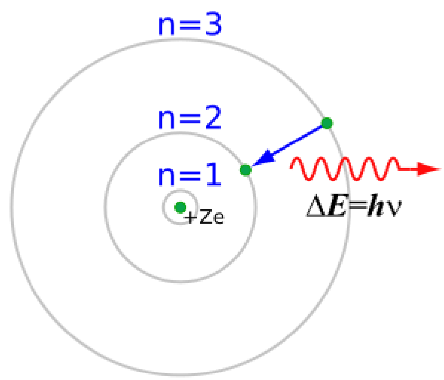
Bohr's Model of the Atom (1913)

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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What High School Chemistry 1 Students Should Know About the Bohr Model

The modern quantum mechanical model (sometimes called the electron cloud model) of the atom evolved from the Bohr model. We now know that electrons don't just follow circular (or elliptical) paths around the nucleus; they move randomly within an "orbital", which is a cloud-like region where there is a higher probability of finding the electron based on how much energy it has.

However, the Bohr model is still helpful even if we now consider it incomplete.

We still believe that:

- The distance between each electron and the nucleus is based on the amount of energy it has.
- The regions where an electron can be found are based on their main or "principal" energy levels. We define these levels the same way that Bohr did: as integer numbers from 1 to 7, which correspond to the periods (rows) of the periodic table of the elements.
- Each energy level can hold a maximum number of electrons. That maximum increases as the energy level gets farther away from the nucleus.
- The outer shell contains the "*valence* electrons". These are the electrons that are available to interact (bond) with other atoms.
  - The first shell can contain a maximum of 2 valence electrons.
  - The other shells can contain a maximum of 8 valence electrons. An atom cannot have more than 8 valence electrons.
- Electrons are added to an atom as follows:
  - The first two electrons in the atom go into the first shell ($n = 1$).
  - After that, new shells are added, starting with $n = 2$, and electrons are added in this order:
    1. A new "valent" shell is added outside of the other shell(s). The first 2 electrons go into this new ("valent") shell.
    2. The next electrons fill up the available spaces in the inner shells. This continues until all of the available spaces in the inner shells are full.
    3. The last 6 electrons go into the new valent shell.
    4. Once the valent shell is full, if there are any more electrons to be added, go back to step #1.

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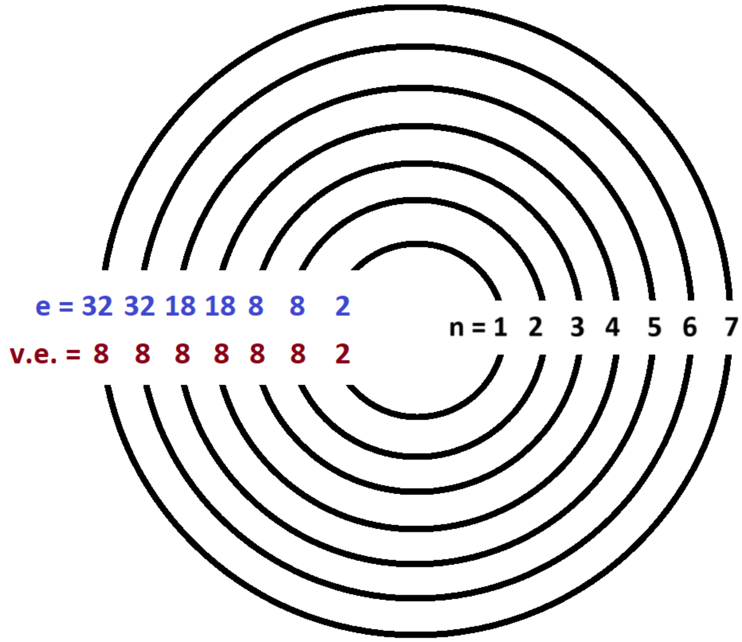
The shells of the Bohr model are arranged like this:

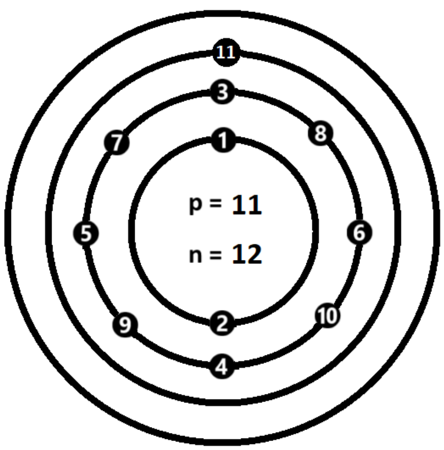


Notice that:

- Shell #1 (n = 1) can hold 2 electrons, both of which are valence electrons.
- Shells #2 and #3 (n = 2 and n = 3) can hold 8 electrons, all of which are valence electrons.
- Shells #4 and #5 can hold 18 electrons, but only 8 are valence electrons.
- Shells #6 and #7 can hold 32 electrons, but only 8 are valence electrons.

For example, a Bohr diagram for $^{23}_{11}Na$, which has 11 electrons, would look like the diagram at the right. Notice that:

- the first 2 electrons are in the inner shell
- the next 8 electrons are in the 2nd shell.
- The last electron is in the 3rd shell.



Remember that the current model of the atom is more complex than the Bohr model represents. However, the Bohr model is useful to understand the idea that electrons exist in different locations around the atom, and those locations are related to the amount of energy that each electron has.

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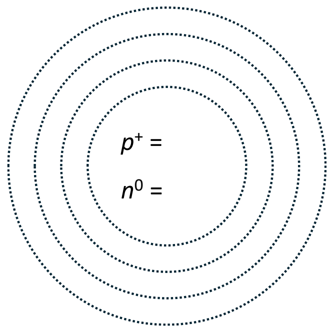
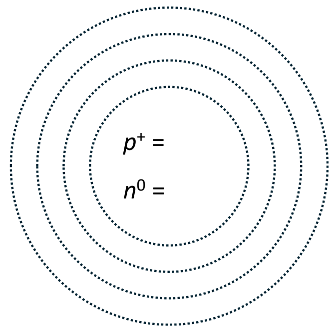
Homework Problems

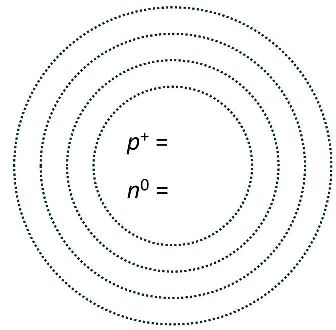
Draw the appropriate number of dots on each ring to represent the electrons in each principal energy level. (*Note: some of the outer rings might not have any electrons, depending on the element.*)



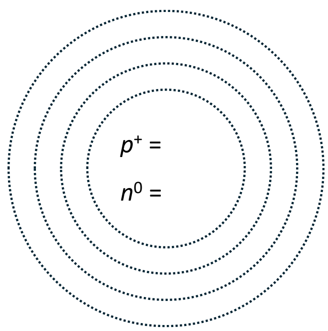
Element: $^{12}_{6}C$
\# e^-:



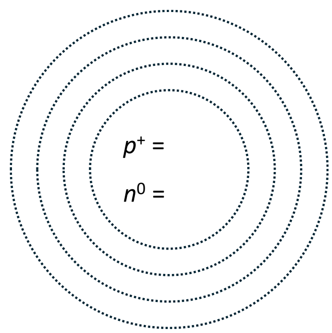
Element: $^{4}_{2}He$
\# e^-:



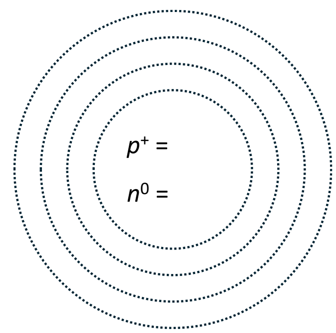
Element: $^{23}_{11}Na$
\# e^-:



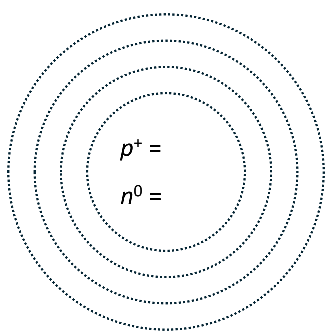
Element: $^{16}_{8}O$
\# e^-:



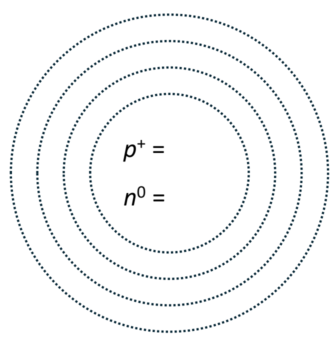
Element: $^{20}_{10}Ne$
\# e^-:



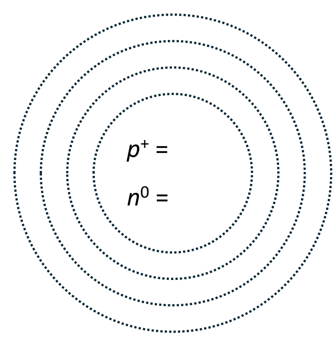
Element: $^{27}_{13}Al$
\# e^-:



Ion: $^{40}_{20}Ca^{2+}$
\# e^-:



Ion: $^{14}_{7}N^{3-}$
\# e^-:



Ion: $^{35}_{17}Cl^-$
\# e^-:

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