Chemistry I
Rydberg and Bohr Equations

Name: Answer Key

Rydberg	more general →	Bohr
$\frac{1}{\lambda} = 1.097 \times 10^7 m^{-1}(\frac{1}{n^2_{final}} - \frac{1}{n^2_{initial}})$		$E = -2.178 \times 10^{-18} J(\frac{Z^2}{n^2})$

Emission/absorption spectra show the energy emitted/absorbed when electrons *change* energy levels (jump between levels). The **Rydberg equation** calculates the wavelength of energy emitted or absorbed by a hydrogen electron *transitioning* from one energy level ($n_{initial}$) to a new energy level (n_{final}). For any single-electron particle (with any Z), the **Bohr equation** calculates the energy of an electron *in* a given energy level.

These two equations differ in three important ways. First, Rydberg calculates wavelength while Bohr calculates energy. Second, Bohr's equation is more general than Rydberg's because it works for single-electron particles other than hydrogen. Third, Bohr's equation doesn't directly show the energy *emitted/absorbed* by transitioning electrons—it only shows the energy of one electron in a particular level

1. Starting from **Bohr's equation**, derive a second equation for the energy emitted/absorbed by an electron **transitioning** from one level ($n_{initial}$) to another (n_{final}) in any single-electron particle. We will do this in class, so take a look at your notes if you're confused.

$$\Delta E = E_{final} - E_{initial}$$

$$\Delta E = -2.178 \times 10^{-18} J \left(\frac{Z^2}{n_f^2}\right) - \left[-2.178 \times 10^{-18} J \left(\frac{Z^2}{n_i^2}\right)\right]$$

$$\Delta E = -2.178 \times 10^{-18} J \left(\frac{Z^2}{n_f^2} - \frac{Z^2}{n_i^2}\right)$$

↳ factored out of above equation

n_f = final E level n_i = initial E level

2. Use the equation you derived in problem #1 to determine the wavelength of light emitted when the electron in Li^{2+} drops from n = 6 to n = 2.

$\Delta E = -2.178 \times 10^{-18} J \left(\frac{3^2}{2^2} - \frac{3^2}{6^2} \right) = -4.356 \times 10^{-18} J$

sign is just direction of E (negative means E emitted)

$4.356 \times 10^{-18} J = 6.626 \times 10^{-34} Js \cdot \nu$

$\nu = 6.574 \times 10^{15} Hz$

$3 \times 10^8 \frac{m}{s} = \lambda (6.574 \times 10^{15} Hz)$

$\lambda = 4.56 \times 10^{-8} m$ 45.6 nm

3. When a hydrogen electron undergoes a quantum leap from n = 6 to n = 2, a photon of light is emitted/absorbed (circle one).

a. Use the Bohr equation you derived in problem #1 to calculate the wavelength of this photon of light. (Hint: the Bohr equation calculates the energy of the photon, and you'll then have to convert that to wavelength.)

$\Delta E = -2.178 \times 10^{-18} J \left(\frac{1}{2^2} - \frac{1}{6^2} \right) = -4.833 \times 10^{-19} J$

$4.833 \times 10^{-19} J = 6.626 \times 10^{-34} Js \cdot \nu \quad \nu = 7.294 \times 10^{14} Hz$

$3 \times 10^8 \frac{m}{s} = \lambda (7.294 \times 10^{14} Hz)$

$\lambda = 4.11 \times 10^{-7} m$ 411 nm

b. Use the Rydberg equation to calculate the wavelength of this same photon of light, and compare your answer to part a. Do you the answers differ? Why or why not?

$\frac{1}{\lambda} = 1.097 \times 10^7 m^{-1} \left(\frac{1}{2^2} - \frac{1}{6^2} \right) = 2.438 \times 10^6 m^{-1}$

now take reciprocal

$\lambda = 4.10 \times 10^{-7} m$

410 nm

very close – just rounding errors... should be the same since we're working with the same transitions in the same atom

4. An electron in an unknown particle transitions from n = 3 to n = 7

   a. Does the electron absorb or emit a photon? How do you know?

   must absorb to be excited
   (move up energy levels)

   b. The wavelength of the photon absorbed or emitted (above) is 2.6×10^{-5} m. What's the energy difference between energy level 3 and energy level 7 in this unknown particle?

$$E = h\nu \qquad c = \lambda\nu \longrightarrow \frac{c}{\lambda} = \nu$$

$$E = \frac{hc}{\lambda}$$

$$E = \frac{6.626 \times 10^{-34}\,\text{Js}\,(3 \times 10^{8}\,\frac{\text{m}}{\text{s}})}{2.6 \times 10^{-5}\,\text{m}} = 7.6 \times 10^{-21}\,\text{J}$$