The Bohr Model

1. What is the Bohr model based on?

The Bohr model is used the hydrogen emission spectrum to create an equation that was built on the idea that electrons could only occupy certain areas of space relative to the nucleus.

2. What is the equation?

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

Which is more frequently used in the form:

$$\Delta E = -2.178 \times 10^{-18}\ \text{J}\ (Z^2)\ \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

n = integer value
Z = atomic number of element
J = Joules

This equation can only be used to determine how much energy is absorbed or released when an electron changes energy orbits.

3. What is the only applicable case to use the Bohr model?

This equation is valid only for atoms with 1 electron, i.e. H, Li^{2+}, C^{5+}, etc.

Another problem with the Bohr model is that it cannot account for bonding; a big issue in chemistry.

4. An electron is excited from ground state to n=3 state in a hydrogen atom. Which of the following statements are true? Correct any false statements.

a. It takes more energy to ionize (remove) the electron from n=3 than from ground state.

False. At n=3 it would be easier as the electrons are farther away from the nucleus (to which they are attracted). Further away something is from what it is attracted to… easier it is to pull them apart.

b. The electron is farther from the nucleus on average in the n=3 state than in ground state.

True

c. The wavelength of light emitted if the electron drops from n=3 to n=2 is shorter that the wavelength of light emitted if the electron falls from n=3 to n=1

False. It would take more energy (therefore a shorter wavelength) to transition from n=3 to n=1

d. The wavelength of light emitted when the electron returns to the ground state from n=3 is the same as the wavelength of light absorbed from n=1 to n=3.

True

5. Does a photon of visible light (λ=400-700nm) have sufficient energy to excite an electron in a hydrogen atom from the n=1 to the n=5 energy

state?

In order to answer this question we are going to have to calculate the minimum energy required to excite the electron from n=1 to n=5. Remember that because of the inverse relationship between E and λ… the λ that we solve for (based on the E value) will pertain to the maximum wavelength possible.

Because we are dealing with hydrogen (which has one electron), we are able to use the Bohr equation:

$$\Delta E = -2.178 \times 10^{-18}\ \text{J}\ (Z^2)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

Where:
$Z = 1$
$n_i = 1$
$n_f = 5$

Plugging in, we get:

$$\Delta E = -2.178 \times 10^{-18}\ \text{J}\ (1)^2\left(\frac{1}{5^2} - \frac{1}{1^2}\right) = 2.09 \times 10^{-18}\ \text{J}$$

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34}\ \text{Js})(2.998 \times 10^{8}\ \text{m/s})}{2.09 \times 10^{-18}\ \text{J}}$$

$$\lambda = 95\ \text{nm}$$

This represents the maximum wavelength of light that would have sufficient amount of energy to excite an electron from n=1 to n=5. So visible light's (400-700nm) range it too great. Visible light would not be able to excite the electron in this way.

6. An excited hydrogen atom emits light with a wavelength of 397.2 nm to reach the energy level for which n=2. In which principal quantum level did the electron begin?

In this problem we have to determine the energy associated with the wavelength of light that was emitted. This will allow us to use the Bohr equation to solve for n_i.

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ Js)(2.998 \times 10^{8}\ m/s)}{397.2 \times 10^{-9}\ m} = 5.00 \times 10^{-19}\ J$$

We can now plug into the Bohr Equation. One *very* important point to note is that the energy here is *emitted*. This means that the electron is releasing this amount of energy upon descent. So when we use the E value, we just calculated, in the Bohr equation – it must be negative.

$$\Delta E = -2.178 \times 10^{-18}\ J\ (Z^2)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

Where:

$\Delta E = -5.00 \times 10^{-19}$ J $\quad n_f = 2$

$Z = 1 \quad n_i = ?$

Plugging in we get:

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

$$\boxed{n_i = 7}$$

7. Consider an electron for a hydrogen atom in an excited state. The maximum wavelength of electromagnetic radiation that can ionize the electron from the H atom is 1460 nm. Determine the initial excited state for the electron.

This is the same procedure as normal. Determine the energy associated with this wavelength of light. Then use this energy to determine the n value using the Bohr equation.

The important thing to realize about this question is what it means to ionize. Ionization is the complete removal of an electron from an atom. This would require energy to be put in. Additionally this means that the final n value would equal infinity, the electron would be infinitely far away from the nucleus if it was removed.

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ Js)(2.998 \times 10^{8}\ m/s)}{1460 \times 10^{-9}\ m} = 1.36 \times 10^{-19} J$$

Now we can use the Bohr equation:

$$\Delta E = -2.178 \times 10^{-18} J\ (Z^2)\ \left(\frac{1}{n^2_f} - \frac{1}{n^2_i}\right)$$

Where:

$\Delta E = 1.36 \times 10^{-19} J$ $\quad$ $n_f = \infty$

$Z = 1$ $\quad$ $n_i = ?$

Plugging in:

$$1.36 \times 10^{-19} J = -2.178 \times 10^{-18} J\ (1)^2 \left(\frac{1}{\infty^2} - \frac{1}{n^2_i}\right)$$

$$\frac{1}{n^2_i} = 0.0625$$

$$n = 4$$