The Bohr Model of Hydrogen Activity

Bohr's model of the atom tried to account for the emission (aka bright line) and absorption (aka dark line) spectra of atoms, along with Planck's theory of energy being quantized in atoms. In the Bohr model, electrons orbited around the nucleus only at specific fixed distances called *energy levels*, symbolized by the lowercase letter *n*. The electrons could only exist at one of the energy levels, but not in between. In fact, based on Bohr's calculations, the electrons could only be found at these specific distances, and never in between. The electrons could even "transition" or move, from level to level, but not in the traditional sense of moving; they didn't actually move like moving a ball off of the floor and setting it on a table. In that case, the ball actually exists at each distance in between the two levels of the floor and the table top. When an electron transitioned from one level to another, it, well, disappeared at one level and reappeared at the other – but was never in between. (I know – that's weird.)

The closer the orbit distance was to the nucleus, the lower the energy; thus, the closest orbit was referred to as the "first energy level", or n=1. The next possible distance away was referred to as energy level 2 (n=2) and so on, out towards an electron being so far away from the nucleus it had essentially left the atom completely, and was described as n=∞. Because energy is released whenever an attractive force is allowed to form, and the nucleus and electron are attracted to each other (coulomb's law), by convention an electron at an infinite distance from the nucleus is defined to have E = 0 kJ/mol, and as an electron moves closer to the nucleus, _the energy of the electron is expressed as a negative number_. But here's the catch – the electron cannot be at just any random distance from the nucleus; it can only exist within the atom at one of the fixed distances away from the nucleus, or orbits. If an electron moved from a higher energy level (further away from the nucleus) to a lower level, energy was released; and, in fact, because these orbits were at specific distances apart from each other, the energy released could be calculated. The energy was released as a photon of light. Because the energy difference between, say level 3 and level 2 was always exactly the same amount of energy, the same photon of light was released every time. And because there are a limited number of possible orbits, there is only set number of energy level differences; that is, there is only a set of specific photons that can be released by any atom. This accounts for the emission spectrum of an atom (rather nicely, actually). The absorption spectrum of an atom lines up exactly with the emission spectrum – the energy of the photons an atom can release is exactly the same as the energy of the photons it can absorb. The dark lines in the absorption spectrum of an element line up exactly with the bright lines of its emission spectrum.

The math…

The wave equation for light:
- $c = \lambda \nu$ (the speed of light = wavelength x frequency)
- c = 3.00×10^8 m/s = 3.00×10^{17} nm/s

The Planck equation: $E = h\nu = \dfrac{hc}{\lambda}$ {energy = Planck's constant x frequency = (Planck's constant x speed of light) / wavelength)}
- h = 6.63×10^{-34} J·s
- allows us to calculate the energy of a single photon of light ($^J/_{photon}$) given its frequency (or wavelength)
- to convert to the energy of a mole of electrons, multiply by Avogadro's number, (6.02×10^{23} photons/mol)

A useful rearrangement of the Planck equation: $\lambda = \dfrac{hc}{E}$
- allows us to directly convert the energy of a photon of light into its corresponding wavelength

The Bohr equation: $E = -R_H(\dfrac{1}{n^2})$ (energy = Rydberg constant x 1/energy level²)
- R_H = 2.18×10^{-18} J (for hydrogen – *hence the subscript H* - other elements would have a different value)
- allows us to calculate the energy of an electron at a specific orbit (energy level)
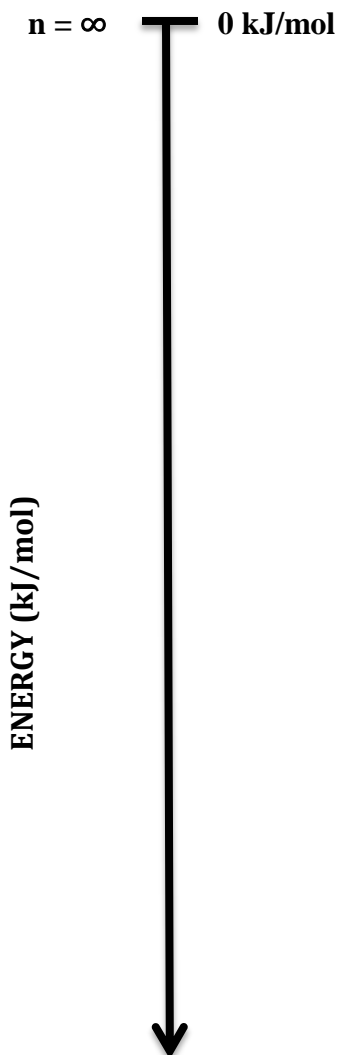
The equation to calculate the energy of a photon released (or absorbed) by an allowable electronic transition (a fancy way of saying an electron is moving from one level to another):

$$\Delta E = R_H(\dfrac{1}{n_i^2} - \dfrac{1}{n_f^2})$$ (where n_i is the initial energy level, and n_f is the final one)

1. Using the Bohr equation, calculate the energy of an electron at each possible energy level ("orbit") in a hydrogen atom.
   Note: To make the numbers more convenient to write, convert your answers from J/photon to kJ/mol by dividing by 1000
   (1kJ = 1000J) then multiplying by Avogadro's number (6.02×10^{23} photons = 1 "mole" of photons).

n (energy level)	Energy (kJ/mol)	n (energy level)	Energy (kJ/mol)
1		6	
2		7	
3		8	
4		9	
5		10	

2. a) When the value for n is getting larger, are you moving further away from the nucleus or closer?

   b) What do you notice about the difference in energy between consecutive energy levels as you move further and further
   away from the nucleus? That is, as you move further away from the nucleus, are the energy levels closer together or
   farther apart?

   c) Sketch the energy for **n = 1 to n = 6** on the graph below, using a scale of 1cm = 100kJ/mol. Indicate the n value on the left
   of the y axis, and the energy amount on the right. The value for an electron leaving an atom completely (moving an
   infinite distance from the nucleus) is included as 0 J.

n = ∞ ⊤ **0 kJ/mol**

ENERGY (kJ/mol)

3. Fill in the tables below as follows.

I. Calculate the ΔE for the transition from a higher energy level to a lower level using the equation on page one. Record the values in the appropriate blank. Leave your answers in J/photon.

HIGHER (INITIAL) ENERGY LEVEL

ΔE (J/photon)	6	5	4	3	2
1					
2					
3					
4					
5					

LOWER (FINAL) ENERGY LEVEL

II. Calculate the corresponding wavelength (in nm) for the photon of light released for each of the electronic transitions in the table above. *Note: the energy the photon possesses is the positive of the ΔE calculated in the table above.*

HIGHER (INITIAL) ENERGY LEVEL

λ (nm)	6	5	4	3	2
1					
2					
3					
4					
5					

LOWER (FINAL) ENERGY LEVEL

In the electromagnetic spectrum, UV light falls in the wavelength range from 10 – 400nm; it is invisible to our eyes, but can cause injury (like sunburns). Visible light is 400 – 760nm; infra-red light is the next longest wavelength, from the far red edge of visible light all the way to around 1 mm ($1x10^6$ nm). We cannot see infra-red light, but we feel it as heat. When you feel heat from a campfire, for example, you are feeling the infra-red photons that are being released.

Which of the three ranges of light, UV, visible, and IR, is the highest energy?

In a hydrogen atom, when an electron "falls" from a higher energy level down to n=1, the photons released are referred to as the "Lyman series". What range of light are they? Can you see them with your eyes?

In a hydrogen atom, when an electron "falls" from a higher energy level down to n=2, the photons released are referred to as the "Balmer series". What range of light are they? Can you see them with your eyes?