Lesson 4.2 hwk answers →

- Each orbit corresponds to a discrete energy - an "energy level". (symbol *n*) Energy levels can be "empty" (i.e. no electron has that energy) or "full" (i.e. some maximum number of electrons have that energy)

- The energy of an electron depends on the size of the orbit, and a smaller orbit means a lower energy. A hydrogen atom with its (sole) electron in $n = 1$ is perfectly stable.

- Electrons usually exist in "the ground state," or in an unexcited state. While it is in the ground state, the electron is not emitting radiation.

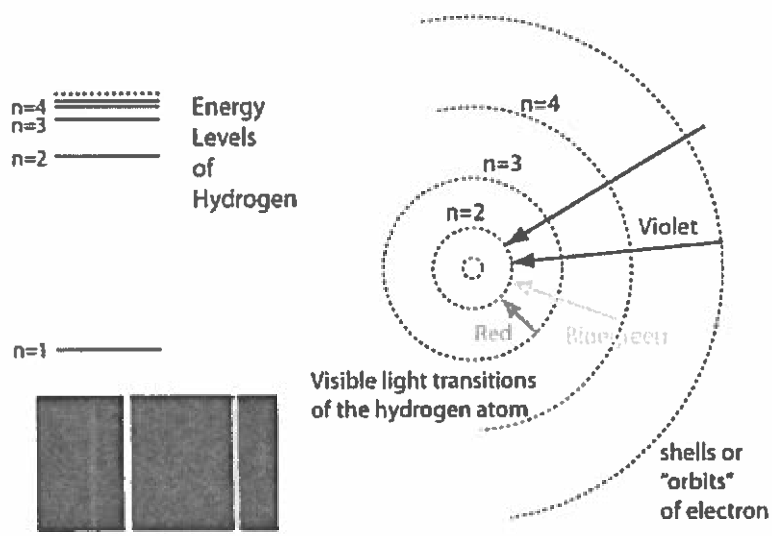
- Electrons in the ground state will absorb energy to "jump up" to higher energy levels, thus becoming "excited" when they do.

- Excited electrons will emit energy to "fall down to" lower energy levels, or back to the ground state.

- Since the energy levels are discrete, the energy that electrons absorb and release is discrete. (This idea was taken from Planck.)

- When excited electrons fall back down to lower energy levels, they emit light of discrete energies and frequencies.

- If the light is visible, you see a photon corresponding to that energy and frequency, following the equation $E = h\nu$. Each color thus corresponds to a line in that element's atomic emission spectrum.



[3]

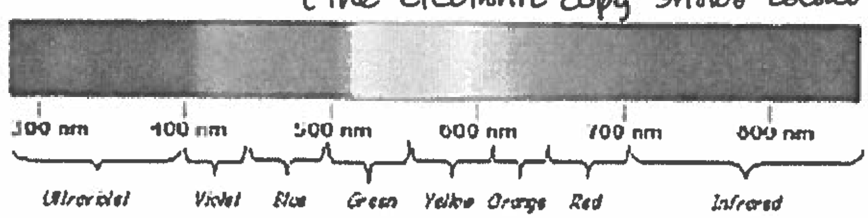
[3] "Modern Atomic Theory." *JAHSChem - Modern Atomic Theory.* N.p., n.d. Web. 14 Dec. 2016.

4.2 Practice and Homework

1. Calculate each quantity and fill in the chart with your answers. See EM spectrum chart on pg 10

Wavelength (m)	Frequency (Hz)	Energy (J)	Electromagnetic Radiation Range
0.001	3×10^{11} Hz	2×10^{-22} J	infrared
4.3×10^{-6} m	7.0×10^{13}	4.6×10^{-20} J	infrared/visible
5.0×10^{-7}	6.0×10^{14} Hz	4.0×10^{-19} J	visible
9.9×10^{-11} m	3.0×10^{18} Hz	2.0×10^{-15}	x-ray
2.5×10^{-14} m	1.2×10^{22}	8.0×10^{-12} J	gamma ray

(the electronic copy shows colors better)



Examine the continuous spectrum above to help you answer question #2.

2. Which color of *visible* light (i.e. ROYGBIV) has...

   a) the shortest wavelength? violet
   b) the longest wavelength? red
   c) the lowest energy? red
   d) the greatest energy? violet

3. The "Balmer series" describes spectral line emissions in a hydrogen atom which correspond to "falls" by excited electrons from higher energy levels to n = 2. For each of these four "falls", calculate the energy (in electron volts) and the frequency of the released radiation. The wavelength of each colored line is in the spectrum graphic on pg. 12.

1) $\lambda = 410\text{ nm} = 4.10 \times 10^{-7}\text{ m}$

$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8}{4.10 \times 10^{-7}} = 7.32 \times 10^{14}\text{ s}^{-1}$

$E = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \cdot 3.00 \times 10^8}{4.10 \times 10^{-7}}$

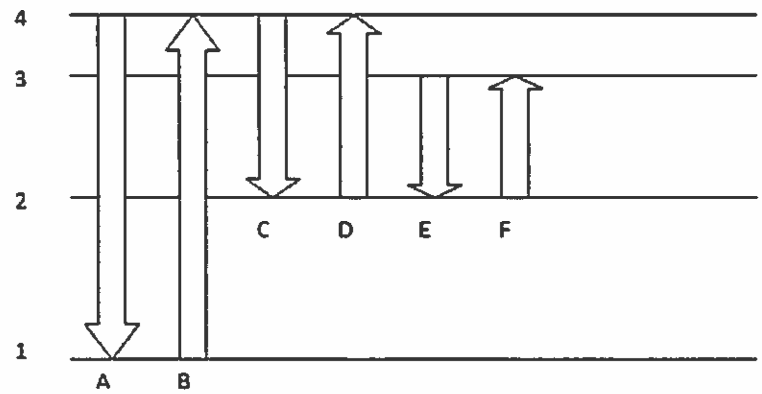
$E = 4.85 \times 10^{-19}\text{ J}$; $\frac{4.85 \times 10^{-19}\text{ J}}{} \Big| \frac{1\text{ eV}}{1.602 \times 10^{-19}\text{ J}} = 3.03\text{ eV}$

In the same way...

2) $\lambda = 4.34 \times 10^{-7}\text{ m}$
$\nu = 6.91 \times 10^{14}\text{ s}^{-1}$
$E = 4.58 \times 10^{-19}\text{ J}$
$= 2.86\text{ eV}$

3) $\lambda = 4.86 \times 10^{-7}\text{ m}$
$\nu = 6.17 \times 10^{14}\text{ s}^{-1}$
$E = 4.09 \times 10^{-19}\text{ J}$
$= 2.55\text{ eV}$

4) $\lambda = 6.56 \times 10^{-7}\text{ m}$
$\nu = 4.57 \times 10^{14}\text{ s}^{-1}$
$E = 3.02 \times 10^{-19}\text{ J}$
$= 1.89\text{ eV}$



4. The figure above represents the four lowest energy levels of an atom. (n = 1 to 4). The six lettered arrows represent changes in the energy level of an electron.

a) Why do these energy levels mean that the atom will show an emission spectrum of discrete lines rather than a continuous spectrum of emitted light?

Each individual energy level represents one specific energy value for the electron in that level - so moving between energy levels causes a change in energy that is also a discrete value.

b) Which three of the lettered energy changes involve absorption of energy by the atom?

B, D, F (absorption of quanta occurs when electrons move up in energy levels)

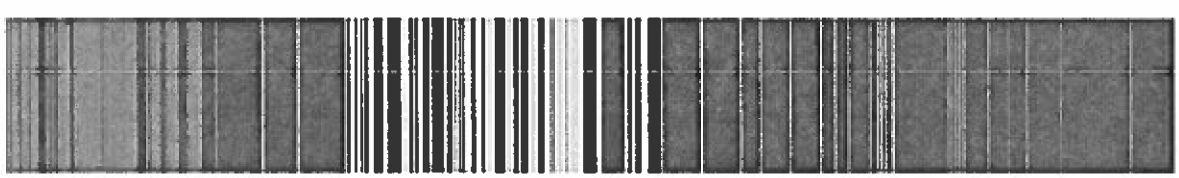
c) Which three of the lettered energy changes involve emission of light energy by the atom?

A, C, E (emission of quanta occurs when electrons move down in energy levels)

Lesson 4.3 and 4.4: Electron Configuration and Quantum Numbers

A. Problems with Bohr's Atomic Model

A scientific model is accepted to be true if it is able to explain phenomena. While Bohr's model does an excellent job of explaining behavior in a one-electron system, such as H, Li^+, Be^+, etc., it cannot successfully predict multi-electron system spectra, such as iron[4]:



The Bohr model cannot predict the intensities of the lines, nor their finer structure - for instance, how can you explain why some lines are thicker or thinner, or why some lines appear to be the same color but have a different placement in the spectrum? Furthermore, Bohr's "stationary states" were purely circular (i.e. in one plane), but atoms are three dimensional.

[4] *Wikipedia.* Wikimedia Foundation, n.d. Web. 14 Dec. 2016.

More relevant for today, the Bohr atomic model also violates the Heisenberg uncertainty principle. This principle states that the more precisely you know the location of a particle (such as an electron in an energy level), the less precisely you know its momentum. Momentum is the quantity of motion that a particle has. In Bohr's model, the position of an electron is an atom is certain, since it's a certain distance from the nucleus, and the energy of that electron is certain, because that distance is related to its energy. This is a contradiction because the velocity of an electron is related to energy: kinetic energy (KE) = $\frac{1}{2} mv^2$. In the Bohr model, you know both position and velocity; according to the uncertainty principle, you can know one, but not the other.

So, we use another atomic model to better explain experimental observations of the atom. We don't completely reject Bohr's statements; that is to say, we cite parts of it in this newer model, and refine the parts we don't.

B. The Quantum Mechanical Model

In 1926, Erwin Schrodinger published mathematical equations describing the probability of finding an electron in a certain position. In other words, you cannot be completely certain where the electrons are. The equations don't define the trajectory of the electron, as the Bohr model does, so concerns about violating the Heisenberg uncertainty principle are moot. These equations are the basis for the quantum mechanical model of the atom. The equations are based on an idea put forth by Louis deBroglie in 1924 stating that electrons in atoms have both wave and particle properties; so, moving matter has a wavelength, which is significant for electrons because of their very small mass. Schrodinger's equations treat electrons like they are waves.

The quantum mechanical model retains Bohr's idea of energy levels, but states that electrons move within an "electron cloud" around the nucleus. Electrons could be anywhere within this electron cloud, but, depending on their energy, they are more likely to be in that corrresponding energy level. The energy levels are areas of higher density within the cloud.

Quantum Numbers

The probable location of an electron is based on four factors (variables in Schrodinger's equations), each represented by a different letter: its principal energy level (n), its sublevel (l), its atomic orbital (m_l), and its spin (m_s). These four values are collectively known as quantum numbers and each has a name: principal energy level, angular momentum number, magnetic quantum number and spin quantum number. There are very specific rules about what these numbers can be. Each electron in an electron cloud has a unique set of "allowable" quantum numbers that fit these rules. Sets of quantum numbers that don't follow these rules are "not allowable", meaning that electron with that set of numbers doesn't exist.

- A principal energy level (n) is one of seven levels inside the electron cloud. The first energy level, n = 1, is the smallest, the least energetic, and the closest to the nucleus. Each level corresponds to a period in the periodic table.

  Allowable values of n must be positive integers greater than 1, i.e. $n = 1$, $n = 2$, $n = 3$, etc.

- A sublevel (l) is a volume of space within an energy level. An energy level can contain several sublevels.