PLI #10
18 June, 2021

1.

1. **Calculate the energy required to excite the hydrogen electron from level n = 1 to n = 2. Also calculate the wavelength of light that must be absorbed by a hydrogen atom in its ground state to reach this excited state (R = 1.097x10^7, h = 6.626x10^{-34} m^{2}kg/s, c=3.00*10^8 m/s)**

   The energy corresponding to allowed orbits for the electron in the hydrogen atom is represented by the formula:

$$E = (-hcR)\left(\frac{1}{n^2}\right)$$

   Therefore, we can consider the change in energy for the electron transition:

$$\Delta E = E_f - E_i = (-hcR)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\Delta E = E_f - E_i = (-6.626 * 10^{-34} J * s)\left(3.00 * 10^8 \frac{m}{s}\right)(1.097 * 10^7 m^{-1})\left(\frac{1}{2^2} - \frac{1}{1^2}\right)$$

$$\Delta E = E_f - E_i = -2.18 * 10^{-18} J\left(\frac{1}{4} - \frac{1}{1}\right) = 1.64 * 10^{-18}\,J$$

   This should make sense because energy must be absorbed (positive value) to excite the electron from n = 1 to n = 2.

   To calculate the wavelength of light,

$$E = h\nu$$

$$c = \lambda\nu \; so \; \nu = \frac{c}{\lambda}$$

$$E = \frac{hc}{\lambda} = \frac{(6.626 * 10^{-34} J * s)\left(3.00 * 10^8 \frac{m}{s}\right)}{\lambda} = 1.64 * 10^{-18} J$$

$$\lambda = \frac{(6.626 * 10^{-34} J * s)\left(3.00 * 10^8 \frac{m}{s}\right)}{1.64 * 10^{-18} J} = 1.21 * 10^{-7}\,m$$

2. **An energy of 3.3*10^{-19}J/atom is required to cause a cesium atom on a metal surface to lose an electron. Calculate the longest possible wavelength of light that can ionize a cesium atom**

   The longest possible wavelength of light will have the minimum energy to ionize a cesium atom. Light with shorter wavelength has greater energy, and thus light with wavelengths shorter than the solution will indeed have enough to energy to ionize the cesium atom.

$E_{photon} = \frac{hc}{\lambda}$

$$\lambda = \frac{hc}{E_{photon}} = \frac{(6.626 * 10^{-34} J * s)\left(3.00 * 10^{8} \frac{m}{s}\right)}{3.3 * 10^{-19}\, J} = 6.02 * 10^{-7}\, m = 602\, nm$$

3. Which set of quantum numbers cannot occur together to specify an orbital or a sub-orbital?

a) n=2, l=1, m_l= -1
b) n=3, l=2, m_l= 0
c) n=3, l=3, m_l= 0 (l values have maximum 0-> n-1)
d) n=4, l=3, m_l=0

4. Find the maximum number of electrons that can have these quantum numbers:

a. **n = 3: 18e-** l = 0 (1 orbital), 1 (3 orbitals), 2 (5 orbitals) =total of 9 orbitals and 18 electrons since 2 electrons per orbital
b. **n = 4, ml = 1 (don't worry about anything past f orbitals): 6 e-**. ml = 1 implies that l must be greater than or equal to 1. Since we don't have to consider anything past f orbitals, l can go up to 3. Thus, these quantum numbers can belong to the 4p, 4d, and 4f orbitals, each containing 2 electrons. to a specific orbital orientation, each orbital contains 2 electrons
c. **n = 4, ms = + 1⁄2: 16 e-** l = 0 (1 orbital), 1 (3 orbitals), 2 (5 orbitals), 3 (7 orbitals)-> total of 16 orbitals-> 32 electrons since 2 electrons per orbital□half have one of the two possible spins (+ 1⁄2) ->16 e-
d. **n = 3, l = 2: 10 e-.** This describes the 5 3d orbitals. Each orbital contains 2 electrons, so a total of 10
e. **n = 2, l = 1: 6 e-**. This describes the 3 2p orbitals. Each orbital contains 2 electrons, so a total of 6

5. Calculate the longest and shortest wavelengths of light emitted by electrons in the hydrogen atom that begin in the n = 6 state and then fall to states with smaller values of n.

The largest transition corresponds to the greatest amount of energy being released in the form of light, which will have the shortest wavelength. Thus, the shortest wavelength can be found by considering the transition from n = 6 to n = 1:

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$$E = -Rhc\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$E = -(1.097 * 10^7 m^{-1})(6.626 * 10^{-34} J * s)\left(3.00 * 10^8 \frac{m}{s}\right)\left(\frac{1}{1^2} - \frac{1}{6^2}\right)$$
$$= -2.12 * 10^{-18}\, J$$

$$E_{photon} = \frac{hc}{\lambda}$$

The energy is negative because energy is being emitted, but to find the wavelength, the negative sign can be dropped since only the magnitude of the energy matters. It also wouldn't make sense to have a negative wavelength.

$$\lambda = \frac{hc}{E_{photon}} = \frac{(6.626 * 10^{-34} J * s)\left(3.00 * 10^8 \frac{m}{s}\right)}{2.12 * 10^{-18}\, J} = 9.4 * 10^{-8}\, m = 94\, nm$$

Using similar logic, one can reason that the smallest transition leads to the smallest amount of energy being released in the form of light, corresponding to the longest wavelength. Thus, consider the transition from n = 6 to n = 5.

$$E = -Rhc\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$E = -(1.097 * 10^7 m^{-1})(6.626 * 10^{-34} J * s)\left(3.00 * 10^8 \frac{m}{s}\right)\left(\frac{1}{5^2} - \frac{1}{6^2}\right)$$
$$= -2.67 * 10^{-20}\, J$$

$$\lambda = \frac{hc}{E_{photon}} = \frac{(6.626 * 10^{-34} J * s)\left(3.00 * 10^8 \frac{m}{s}\right)}{2.67 * 10^{-20}\, J} = 7.46 * 10^{-6}\, m = 7.46\, \mu m$$

An excited hydrogen atom emits light with a frequency of 1.141×10^{14} Hz to reach the energy level for which n = 4. In what principal quantum level did the electron begin?

Using the frequency, calculate the energy associated with the emitted light.

$E = h\nu = (6.626 * 10^{-34} J * s)\ (1.141 * 10^{14}\, s^{-1}) = 7.56 * 10^{-20}\, J$

$E = -Rhc\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$

Since light is being emitted, the electron must have a principal quantum level greater than 4, and the energy should be negative.

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$-7.56*10\ J = -(1.097*10^{7}\ m^{-1})\ (6.626*10^{-34}\ J*s)(3.00*10^{8}\ \mathrm{m/s})\left(\frac{1}{4^2} - \frac{1}{n_i^2}\right)$

$0.03467 = \left(\frac{1}{16} - \frac{1}{n_i^2}\right)$

$\frac{1}{n_i^2} = \left(\frac{1}{16} - 0.03467\right)$

$n^2 = \frac{1}{0.02782968}$

$n_i = 6$

The electron transitioned from n = 6 to n = 4.