**Experiment 19: Spectra**

Name:

**Data & Results Table**

 **Part 1 – Flame Tests**

|  |  |
| --- | --- |
| Positive ion tested | Color of flame |
| Na+ |  |
| K+ |  |
| Ca2+ |  |
| Sr2+ |  |
| Ba2+ |  |
| Cu2+ |  |

|  |  |  |
| --- | --- | --- |
| Test solution | Color of flame | Which positive ion is present? |
| Purified water |  |  |
| Tap water |  |  |
| A raisin |  |  |
| liquid soap |  |  |
| Unknown number: |  |  |
| Beverage |  |  |
| Banana |  |  |
| A solution of Tums |  |  |
| Vitamin pill solution |  |  |

**Part 2 – Spectroscopic Examination**

Use crayons or write the name of the color observed to record the spectrum that you see in each case.

Daylight or a regular light bulb, or other

Hydrogen

Mercury

Helium

Sodium or other lamp.

**Questions**

1. Fill in the following table regarding the hydrogen spectrum. The wavelengths for the lines represent the visible region of the electromagnetic spectrum for hydrogen are given below. To complete the chart, you will need to find the ninitial values. Show the calculation for the n initial (ni)l values using

$∆E=2.179x10^{-18}J \left(\frac{1}{n\_{f}^{2}}-\frac{1}{n\_{i}^{2}}\right) $and. $∆E=\frac{hc}{λ}$ . You can also use $\frac{1}{λ}=1.096 8x10^{7}m^{-1}$ to solve for the wavelength in other problems.

Remember that the negative sign in this equation gives direction to the energy. We recognize that these are emissions and use the absolute value for the energies in the calculations.

**SHOW THE CALCULATION FOR RED DATA HERE**

(put the rest in your lab book)

|  |
| --- |
| Balmer Series |
| Wavelength | Energy | nf | ni |
| Red 656.4 nm |  |  |  |
| Turquoise486.3 |  |  |  |
| dark blue434.2 |  |  |  |
| Purple410.3 |  |  |  |

1. The lines of the Lyman series of hydrogen cannot be observed by the naked eye. The transitions are in the ultraviolet region or higher of the EM spectrum. The wavelengths are 93.8 nm, 95.0 nm, 97.3 nm, 102.6 nm, and 121.6 nm.
2. Determine the energy of each wavelength, nf, and ni

**SHOW THE CALCULATION FOR λ= 93.8 nm DATA HERE**

(put the rest in your lab book)

|  |
| --- |
| Lyman Series |
| Wavelength | Energy | nf | ni |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |

1. A similar series of emissions exists in the infrared region of the hydrogen spectrum. This is the Paschen series. The wavelengths are: 1005 nm, 1094 nm, 1282 nm, and 1875 nm.
2. Determine the energies of these transitions, show your work below, and put your final answers in the chart.
3. To fill in the chart, you will need to calculate, ni, nf, and ∆E,

**SHOW THE CALCULATION FOR** λ**= 1,282 nm DATA HERE**

(put the rest in your lab book)

|  |
| --- |
| Paschen Series |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |
|  |  |  |  |

1. What is the correlation between the difference between ni and nf, ∆E, and λ?
2. For a hydrogen atom, only the series lines from which nf=2 contains frequencies in the visible region. What is the value (**in theory**) for nf for which there are frequencies in the visible region for the H-like ions He+, Li2+, and Be3+? (in reality, these wavelenths might not be in the visible range)
3. The ground state energy of the electron in a He + is -54.38eV. Energies such as these are often measured in electron-volts (eV) and converted to joules or kilojoules.
4. Calculate the ∆E for the transition ni=5→ nf=2 in eV and joules.
5. Draw an arrow showing the transition of an electron in n=5 to n= 2.
6. Calculate the wavelength and frequency of the light that is emitted during the transition ni=5→ nf=2.
7. What region of the electromagnetic spectrum does this line occur? This is not a hydrogen atom, so it will not exhibit the same type of transitions that hydrogen undergoes, so it might not be the visible region. You might have to look up the region.

1/λ =RZ2(1/nf2 – 1/ni2), 1.6022x10-19J = 1 eV

1

2

3

4

5

2

1. The energy diagram for hydrogen “like” ion (an ion with one electron) is given below.

|  |  |
| --- | --- |
| 0 |  |
|  |  |
|  |  |
|  |  |
|  |  |
|  |  |
|  |  |
|  |  |
| -3.4 eV |  |
| -4.89eV |  |
| -7.65 eV |  |
| -13.60 eV |  |
| -30.59eV |  |
| -122.36eV |  |

Nucleus of atom

1. What is happening to the energy levels as the value of the energy released gets smaller?
2. Does this mean it takes more or less energy to remove the electron from the nucleus?
3. Using only the energy level diagram, how much energy is required to remove the electron from its nucleus, so that it is gone from the atom? You can also calculate the energy that is released when the electron is added when an electron starts at n= ∞ and ends at the n=1 position of the energy diagram.
4. Identify the ion by determining its atomic number Z. [Hint 1: Egs H = 13.60eV, and you know the Egs for He. Hint 2: don’t be fooled by the numbers, sometimes a cigar is just a cigar.]
5. On the diagram draw an arrow to indicate the transition n= 3 to n=1
6. Calculate the frequency and wavelength of the radiation emitted for the transition n= 3→n=1..
7. At what region of the electromagnetic spectrum does the line occur?
8. Scientists in the 19th and early 20th century were unable to explain why no electrons are emitted when the frequency of light shined on a metal surface in the below minimum threshold value, although light of the higher frequency causes the emission of electrons. They were also unable to understand the observation that the greater frequency light is, above the threshold frequency, the greater the speed, and therefore, the kinetic energy of the emitted electrons. Einstein first successfully explained the photoelectric effect in 1905. He proposed that there was a minimum amount of energy necessary to eject electron from a metal surface. This minimum energy (the work function of the metal) is a characteristic property of the metal. If a photon of the light striking the metal has energy less than the work function of the metal, no electrons are emitted. If a photon of the light used has energy greater than the work function of the metal, the kinetic energy of the emitted electron equals the difference in the energy between energy of the photon and work function of the metal.
9. If a yellow light of wave length 598.0 nm is shined on a potassium surface, the emitted electrons have a kinetic energy of

5.77 X10 —20J/electron. Calculate the value of the work function (threshold energy) and the threshold frequency for potassium.

1. If UV light of wavelength 253.7 nm is shined on a potassium surface what will the speed of emitted electron be?