EXPERIMENT A10: LINE SPECTRUM

Learning Outcomes

Upon completion of this lab, the student will be able to:

1) Examine the line spectrum of the hydrogen atom.
2) Calculate the frequency and energy of the electronic transition corresponding to each observed line in the spectrum.

Introduction

Much of the knowledge of atomic structure is a result of spectroscopy. Spectroscopy is the analysis of light emitted or absorbed by substances. Light is said to have a "dual nature".

The dual nature of light implies that light has both particle-like and wave-like properties. The particle-like property of light is inferred from it consisting of packets (or quanta) of energy (E) called photons. The wave-like property of light leads it to have a characteristic frequency (ν) and wavelength (λ). The energy of the photon is proportional to its frequency.

The relationship between these aspects is described by the following equation:

$$E = h\nu = \frac{hc}{\lambda}$$

In the above equation: "h" is Planck's constant (6.626×10^{-34} Js) and "c" is the speed of light (3.00×10^{8} m/s).

All light moves through space with the same speed, but its effect on matter depends on its energy. Light in the visible and ultraviolet regions of the electromagnetic spectrum is associated with changes in the energy of electrons in atoms, molecules, or ions. This can lead to various observed effects: neon signs, street lamps, black lights etc.

The lowest possible energy state of an atom is called the ground state. In the ground state all of the electrons are in their lowest possible energy levels, as dictated by the Aufbau principle. When an atom absorbs energy, by absorbing a photon of light, an electron enters an excited state, jumping to a higher energy level. Due to quantization, the energy of the light absorbed by an atom is equal to the difference in the energy of the ground state and the excited state of the electron.

Atoms in their excited states are inherently unstable and must lose their excess energy to return to lower energy states. The amount of energy lost is equal to the difference in energy of the excited and the ground states of the atom.

The amount of energy lost when an electron moves from a higher energy level to a lower energy level is quantized. The magnitude of the emitted energy (or light) is directly proportional to its frequency (and inversely proportional to the wavelength) of the radiation. A plot of emitted radiation as a function of the wavelength of the light is referred to as the **Emission Spectrum**. The emission spectrum of an atom is a characteristic feature of that atom. Due to the quantization of the emitted photon, the emission spectrum consists of discrete lines.

The difference in energy between the energy levels relates to the wavelength of the emitted line according to the following formula:

$$\Delta E = E_{final} - E_{initial} = -2.179 \times 10^{-18} J\left(\frac{1}{n^2_{final}} - \frac{1}{n^2_{initial}}\right) = \frac{hc}{\lambda}$$

In the above equation: "h" is Planck's constant (6.626×10^{-34} Js) and "c" is the speed of light (3.00×10^{8} m/s).

Experimental Design

In this experiment the emission spectrum of hydrogen atom will be studied. The wavelengths of the spectral lines will be used to determine the specific electronic transitions. It should be noted that since the observed wavelengths are all in the visible region of the electromagnetic spectrum, the final energy level of the transitioning electrons will be $n_f = 2$.

Reagents and Supplies

Source of H-atoms, He-atoms, Xe-atoms, spectroscope

(See posted Material Safety Data Sheets)

Procedure

1. Observe the line spectrum of the H-atom through the spectroscope.

2. Record the exact wavelengths of each observed line in the spectrum.

3. Observe the line spectra of He and Xe (if available).

Data Table

Color of line	**Wavelength (nm)**

Data Analysis

1. Using the Bohr's equation, given below, for the allowed energy levels of hydrogen atom, calculate the energy of the first six energy levels. The energies of n = 1, and n = 2 are already calculated and shown in the table. [NOTE: Express all the answers in the same power of 10 (i.e. 10^{-19})].

   Bohr's Equation: $E_n = -\frac{2.179 \times 10^{-18}}{n^2} Joules$

Value of "n"	Energy in J
6	
5	
4	
3	
2	-5.44×10^{-19}
1	-21.79×10^{-19}

2. Complete the table given below by calculating the energy of a photon in Joules and the wavelength in nanometers.

   Energy of the photon = $\Delta E = E_{final} - E_{initial}$

   NOTE: The energy of each energy level, E_n, was already calculated in Question 1 above. These values may be used to calculate ΔE.

Transition($n_{initial}$ to n_{final})	Photon Energy ($\times 10^{-19}$ J) ΔE	Wavelength (nm)
6 ➔ 2		
6 ➔ 1		
5 ➔ 4		
5 ➔ 3		
5 ➔ 2		
5 ➔ 1		
4 ➔ 3		
4 ➔ 2		
4 ➔ 1		
3 ➔ 2		
3 ➔ 1		
2 ➔ 1	**−16.34**	**121.6**

3. Compare your experimental wavelengths with the calculated wavelengths using the table from question 2. Determine the transitions corresponding to the data obtained.

Color of line (page 3 Data)	**Experimental Wavelength (nm) (page 3 Data)**	**Corresponding Calculated Wavelength (nm) (Question 2)**	**Corresponding Transition (Question 2)**

4. Construct an energy level diagram to scale for the allowed energies of the electron in the hydrogen atom. Use the results from question 1 above. On the energy level diagram:
    a. Draw and label the n = 1 and n = 6 energy levels. Give the value of n **and** the energy for each level.
    b. Indicate which energy level is the ground state.
    c. Indicate which transitions correspond to the emission lines observed for hydrogen. Draw the transitions as down arrows in the diagram and label each transition with the wavelength **and** color that is observed.