Spectral Analysis (E11)

Objectives

- Observe and discuss the light spectra coming from various sources.
- Analyze the spectrum of hydrogen gas using Bohr model.

Theory

A. Electromagnetic waves.

The electromagnetic waves can have various wavelength, intensity and polarization. The range of observed wavelengths is enormous (more than twenty orders of magnitude). Radio waves, cellular phone waves, microwaves, infrared radiation, visible light, ultraviolet, X-rays, Gamma rays, cosmic rays are all electromagnetic waves. All these waves propagate in vacuum with the same speed c, but have different values of the wavelength λ (or the frequency f, which is related to the wavelength: $\lambda = c/f$). The interaction of electromagnetic waves with objects varies widely depending on the wavelength. Some waves are harmless for humans (e.g., radio waves, visible light), whereas some are quite dangerous, e.g., gamma rays, X-Rays. The visible light represents a relatively narrow range of wavelengths from $\lambda \cong 380$ nm to $\lambda \cong 750$ nm.

B. Bohr model of the atom.

The Bohr model of a single-electron atom is the first quantum model explaining stability and basic properties of hydrogen atoms. Since Rutherford's experiment with scattering of alpha particles, scientists knew that each atom has a positively charged nucleus and one or more electrons moving around the nucleus. This was a major problem, because according to classical electromagnetism, electrons moving in circular or elliptical orbits accelerate and should constantly emit electromagnetic waves losing energy. Therefore, according to classical physics all atoms should be unstable, which is obviously not true. The Bohr model was the first model involving quantization of angular momentum, energy and orbit radius. The Bohr model has its significant limitations. It does not work for atoms with multiple electrons. Even for hydrogen atoms, quantum mechanics offers now a more accurate model. We will limit our discussion here to the Bohr model.

The Bohr model assumes that electron can only occupy certain, stable orbits. These stable orbits have values of the angular momentum L equal to:

$$L = n\frac{h}{2\pi}, \quad \text{where } h \text{ is Planck's constant and } n = 1, 2, 3, \ldots \tag{1}$$

We say that the angular momentum is <u>quantized</u>, i.e., it can only have discrete values and no values in between these stable orbits. The stable orbits are labeled with the positive integer number n. We say that an electron is in "n"-th orbit, i.e., its angular momentum is equal to:

$$L_n = n\frac{h}{2\pi}.$$

It is not very difficult to show that quantization of angular momentum leads to quantized values of electron's energy. The electron's energy at orbit "n" is equal to:

$$E_n = -\frac{2\pi^2 e^4 k^2 m_e}{h^2}\frac{1}{n^2} = -13.6eV\frac{1}{n^2} = -\frac{13.6eV}{n^2} \tag{2}$$

where: $1\ eV = 1.602*10^{-19}$ J

An electron can occupy any stable orbit. There is an infinite number of available orbits between $n = 1$ and $n = \infty$. The electron can move from one stable orbit to another during absorption or emission of electromagnetic wave. During absorption of a photon, electron jumps to an orbit with a larger energy, i.e., with a larger quantum number n. All energy levels E_n for electron are negative, so for a larger n, the energy becomes less negative, i.e., larger than that for a smaller n. The opposite happens during emission of a photon. Electron jumps to an orbit with a lower n value, loosing (emitting) energy in the form of electromagnetic wave.

Since in the Bohr model electron can only jump between stable, discrete orbits, the emitted photons must have specific, discrete values of energy. Therefore, we do not observe continuous spectrum of photons emitted from a hydrogen atom, but discrete, narrow lines with diverse colors. The energy of the emitted photons can be calculated from equation (2). If an electron jumps from orbit with quantum number n_i to orbit n_f, then the energy of the emitted (or absorbed) photon is equal to

$$E_{photon} = -13.6eV\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \tag{3}$$

The energy of photon is related to its wavelength λ.

$$E_{photon} = hf = \frac{hc}{\lambda} = -13.6eV\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \tag{4}$$

$$\frac{1}{\lambda} = -R\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \tag{5}$$

where $R = 1.097*10^7$ m^{-1} = 0.01097 nm^{-1} is called the <u>Rydberg constant</u>.

Here is an illustration of possible transitions between various electron orbits in hydrogen atom.
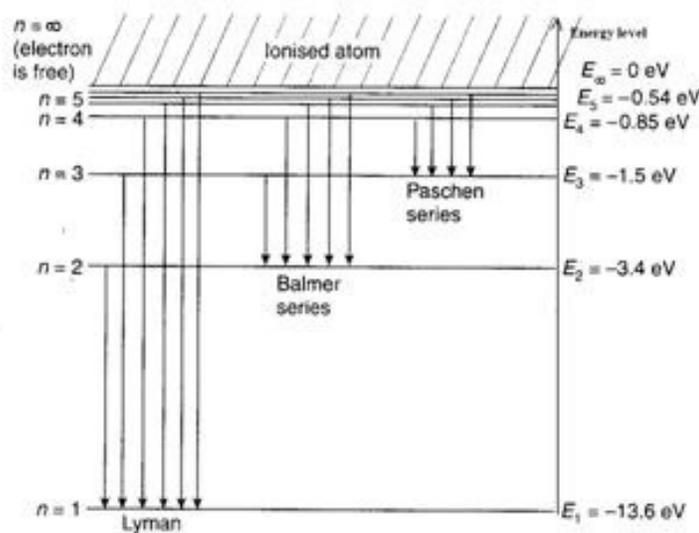
The transitions are arranged in groups called series. All transitions from orbits with the quantum number $n_i \geq 2$ to the first orbit ($n_f = 1$) belong to Lyman series. Transitions from orbits $n_i \geq 3$ to the second orbit ($n_f = 2$) belong to Balmer series. Transitions from orbits $n_i \geq 4$ to the third orbit ($n_f = 3$) belong to Paschen series and so on. Lyman series is in ultraviolet range of the electromagnetic wave spectrum, part of the Balmer series in in the visible range and will be subject of this experiment. Paschen series belongs to infrared radiation.

For Balmer series ($n_f = 2$), the Bohr model predicts emission (or absorption) of electromagnetic waves with the following wavelengths:

$$\frac{1}{\lambda} = -R\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = -R\left(\frac{1}{n_i^2} - \frac{1}{2^2}\right) \qquad \text{where: } i = 3, 4, 5... \tag{6}$$

$$\frac{1}{\lambda} = -R\frac{4 - n_i^2}{4n_i^2} \quad \Rightarrow \quad \lambda = \frac{1}{R} \times \frac{4n_i^2}{n_i^2 - 4} \quad \text{or} \quad R = \frac{1}{\lambda} \times \frac{4n_i^2}{n_i^2 - 4} \tag{7}$$

C. Other sources of light.

Hot metals produce continuous spectrum of light. Metals also contain atoms, but in addition to atomic discrete lines, there is almost infinite number of possible collective modes like vibrations, which leads to a continuum spectrum of light. Example of hot metals: the tungsten filament inside incandescent bulbs.

The light produced by **mercury lamps** seems white but is a combination of several lines. Four of these lines are in the visible spectrum of light and one is in the near ultraviolet.

Lasers. Lasers always emit a single wavelength (a single color) of light. The most common lasers produce red or green light.

D. Absorption spectra.

If light from a source with continuous spectrum (e.g., Sun) goes through a low-density gas, then photons with frequencies corresponding to allowed transitions would be absorbed and absent from the continuous spectrum. When analyzed using a spectrometer, the spectrum would have black lines indicating absorbed photons. Since patterns of atomic spectra are different for each element or compound, this is a method of remote identification of elements or compounds. It is one of the main tools in astrophysics.

Equipment setup.

To turn on the spectrometer, briefly <u>press and hold</u> the ON/OFF button on the side. <u>When all three status LEDs blink in sequence</u>, release the ON/OFF button.
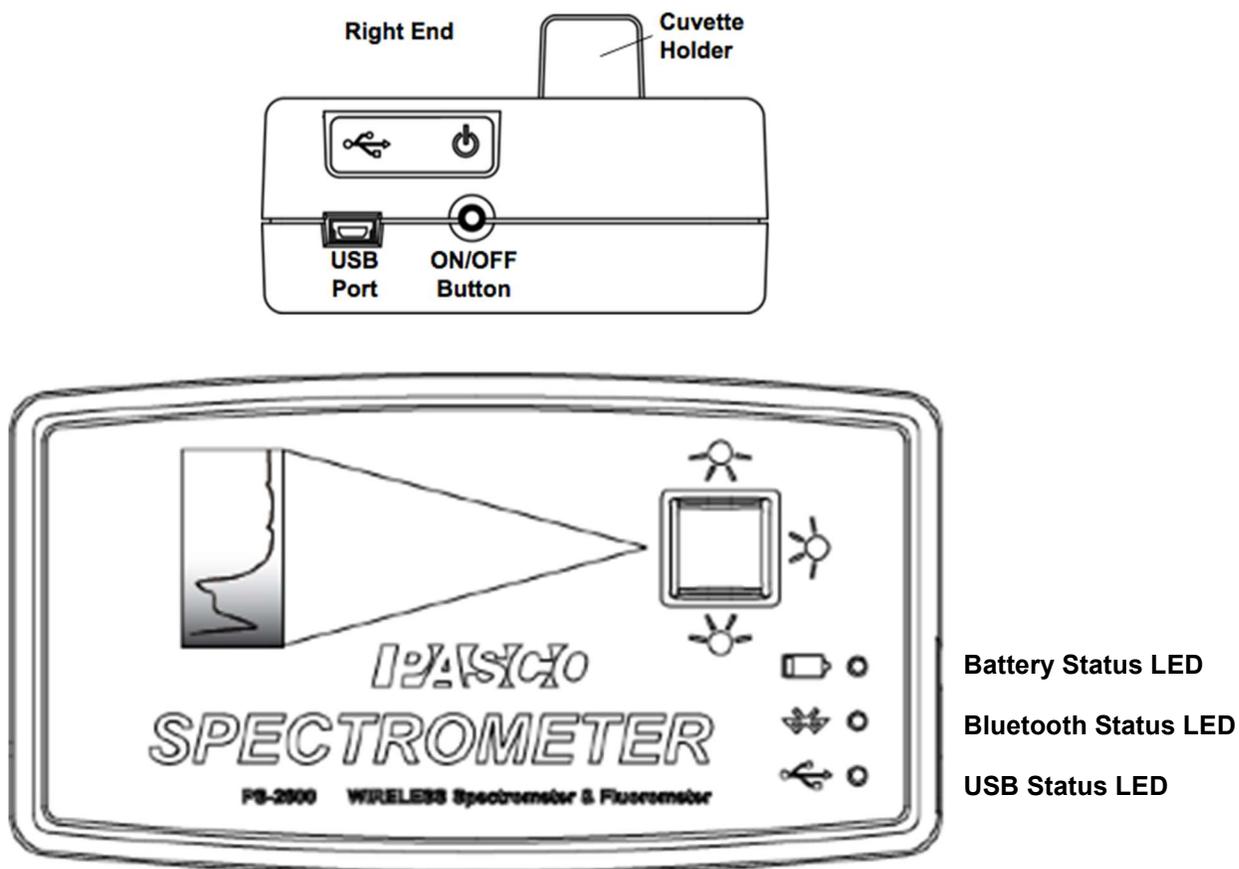
The USB Status LED should continuously shine green. (NOTE: The Bluetooth Status LED will slowly blink blue). The Battery Status LED may also blink to indicate that the battery is charging through the USB cable. We are going to use the USB connection, not the Bluetooth wireless connection between the spectrometer and the computer. The accuracy of our spectrometer (specified by the manufacturer) is approximately equal to ±2 nm.

Software tutorial.

Open the "Spectrometry" application.

We will discuss below only the key features of the "Spectroscopy" application.

The top menu toolbar shows main tools. We use the "**Analyze Light**" option.

From the top menu, you are going to use the following tools:

Analyze a graph of Intensity vs. Wavelength

Analyze a Solution in terms of Intensity vs Wavelength.

Take Journal Snapshot

Show Journal Snapshots

Connection Status

Error

OR

OR

Connected

The Spectroscopy application offers the following graph tools:

Record

Start recording data.

OR

OR

Stop recording data.

Add Coordinate Tool

Add Coordinates tool to select points on the graph.

Note: For additional Solution Analysis with Concentration and Time experiments, an Analysis wavelength MUST be selected.

Show Comparison Mode

Compare spectra of multiple light sources.

Show Color Spectrum Background

Show Color Spectrum Background.

Reference Spectra

Add Reference Lines. Lines include: Hydrogen, Helium, Argon, Xenon, Mercury Sodium, Neon from NIST (National Institute of Standards and Technology) Atomic Spectra Database.

Previous reference

Next reference

The "Reference Spectra" tool is particularly useful. You can compare observed spectrum with reference lines for many elements and see which one matches the observed spectrum.

The "Light Analysis Tools" are located on the left-hand side, and they can be revealed or hidden. There are three "Light Analysis Tools" available:

Integration Time

Pick the best integration time

Integration Time is analogous to the shutter speed of a Camera. With a Higher Integration time, the spectrometer is more sensitive to less intense light.

The "Auto Set" Button will automatically adjust the Integration time to maximize the current sample spectrum.

Number of Scans to Average

Number of Scans to Average

This property specifies the number of discrete acquisitions that are collected and averaged before displaying and recording a spectrum or data point.

A higher the value results in a better the signal-to-noise ratio.

Smoothing

Smoothing

Average groups of adjacent data points.

Whenever you change the light source or you change the position of the light probe, click on the "Integration Time **Auto Set**" button to get the correct exposure, i.e., to get the correct amplitude of the spectrum.

Make sure to complete the following tasks:

You must submit the answers to the prelaboratory questions online.   (3.5 points)

1. *The three graphs from Activities 1÷3.*   (3 × 1 point = 3 points)
    (Title and <u>write your name</u> and those of your partners on the graph.)

2. Your completed Data Sheets.   (3.5 points)

3. Return the completed lab report to your lab TA.