Lab 03: Atomic Emission Spectra

INTRODUCTION

Goals

- Understanding the nature of light, electron wave, and their interaction.
- Describing how the absorption or emission of light is related to changes in energy states of an atom's electron waves.
- Explaining the quantum nature of light and electrons from atomic emission spectra.

Background

Visible light is a form of electromagnetic radiation that propagates through space and time as a wave. The wave consists of electric and magnetic fields that oscillate perpendicular (at 90° angles) to each other and the direction the wave is traveling, as show in Figure 1. The oscillations can be mathematically modeled as a cosine or sine wave dependent on the position and space and time. The **intensity** (brightness) of the light is related to the **amplitude** (A) of the wave. The **color** of the light is related to its **wavelength** (λ), and visible light typically has a wavelength in the range λ = 380 nm to 780 nm.

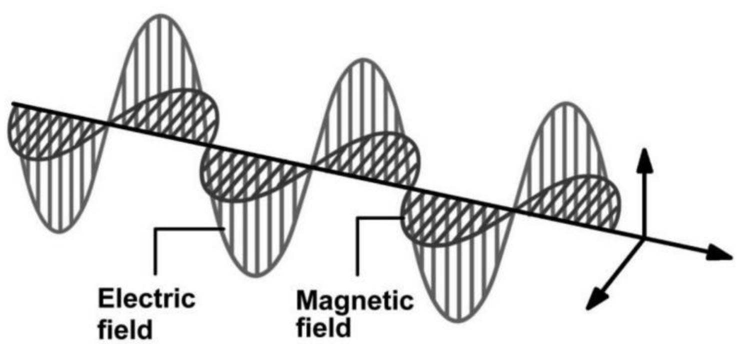

Figure 1. Diagram of an electromagnetic wave. The amplitude A of the wave is the maximum value of the electric field at the peak of an oscillation, the wavelength λ is the distance between two peaks of the wave in space, and the frequency ν is the number of oscillations that pass a fixed point in space per second, measured in Hertz (Hz), with one Hertz equal to one inverse second, $1\ Hz = 1\ s^{-1}$.

The wavelength of light (λ) is related to its frequency (ν) and velocity. Electromagnetic radiation in empty space or air propagates at a constant velocity, known as the speed of light, $\mathrm{c} = 2.9979 \times 10^{8}\ \frac{\mathrm{m}}{\mathrm{s}}$. The relationship between the three parameters is described by the following equation:

$$\nu = \frac{c}{\lambda} \tag{1}$$

Light can interact with matter by transferring energy to and from the electron clouds of the matter. When light is **absorbed** by matter, energy is transferred from light to the matter upon many (on the order of 10^5) oscillations of the electric field. Light is **emitted** when energy from matter (the oscillations of the electron waves) is transferred into a propagating electromagnetic wave. A simplified diagram illustrating the principle of light-matter interaction is shown in Figure 2. *You should be able to draw this diagram on your own!*

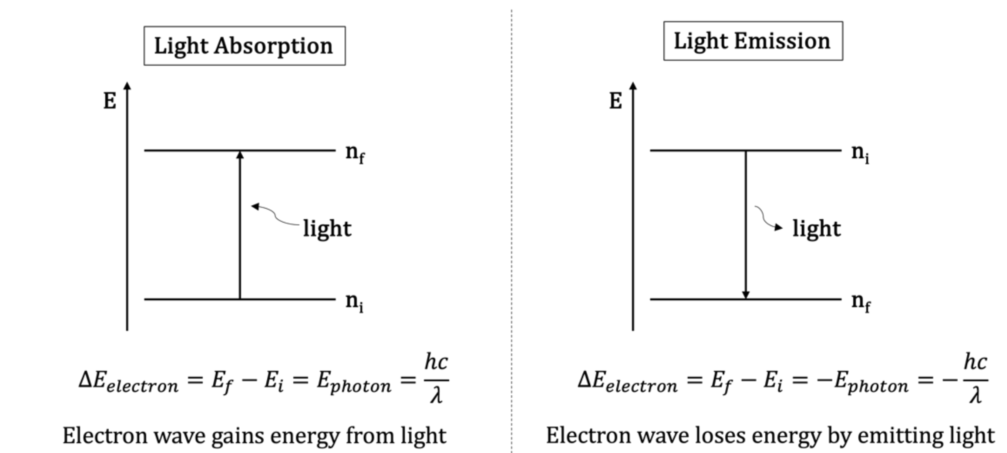

Figure 2. A one-dimensional energy diagram showing the interaction between a matter's electron wave and light. n_i and n_f represent the energy state of the electron wave before and after the interaction, respectively. Note E_{photon} is always a positive number (you cannot have a negative amount of energy!)

The transfer of energy by absorption or emission of light only happens at a specific frequency, and the amount of energy can be transferred is known as the **photon** energy E_{photon}

$$E_{photon} = h\nu = \frac{hc}{\lambda} \tag{2}$$

where $h = 6.626 \times 10^{-34}$ J $\cdot$ s is a constant known as the Planck's constant. Each time light interacts with an electron wave, exactly one photon of energy is transferred. For this energy transfer to happen, the frequency of the absorbed or emitted light must be equal to or, in other words, in **resonance** with the change in frequency for the matter's electron wave.

In this experiment, we will observe the **emission spectrum** of different samples of atomic gases – hydrogen gas, $H_2(g)$, helium gas, $He(g)$, and mercury, $Hg(g)$. The electrons in a sample of atomic gas are typically in their lowest energy state, also known as the **ground state**. When we apply a high voltage of the discharge lamp to the atomic gas, the electrons gain energy and are excited into a higher energy state, described as being in an **excited state**. The electrons in excited states are less stable and can relax back to a lower energy state by emitting light. *Can you draw an energy diagram representing the scenario depicted here?*

Recall that the energy of the electron waves in the hydrogen atom is given by:

$$E_n = -(2.1799 \text{ aJ})\frac{1}{n^2} \tag{3}$$

where n is the principal quantum number, a positive integer that describes the wave-like behavior of the electron (n is equal to the number of "loops" in the electron cloud). When the electron is in its excited state (n > 1), it can lose energy by transitioning to a lower energy state with a lower value of n, and in the meantime emit a photon of energy of light. If we apply equation (3) to the emission diagram shown in *Figure 2*, we get:

$$\Delta E_{electron} = E_f - E_i = (-2.1799 \text{ aJ})\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = -E_{photon} = -\frac{hc}{\lambda} \tag{4}$$

By removing the negative signs on both sides, we get:

$$(2.1799 \text{ aJ})\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = \frac{hc}{\lambda} \tag{5}$$

Equation (5) suggests that the wavelength of the emitted light (λ) depends on the final and initial states of the electron wave, n_f and n_i. Therefore, we can expect the emission spectrum for an atom will consist of **discrete** lines that correspond to these energy transitions, as n_f and n_i are integers, not continuous numbers. An example spectrum for the emission spectrum of hydrogen and a diagram of the corresponding energy transitions are shown in Figure 3.

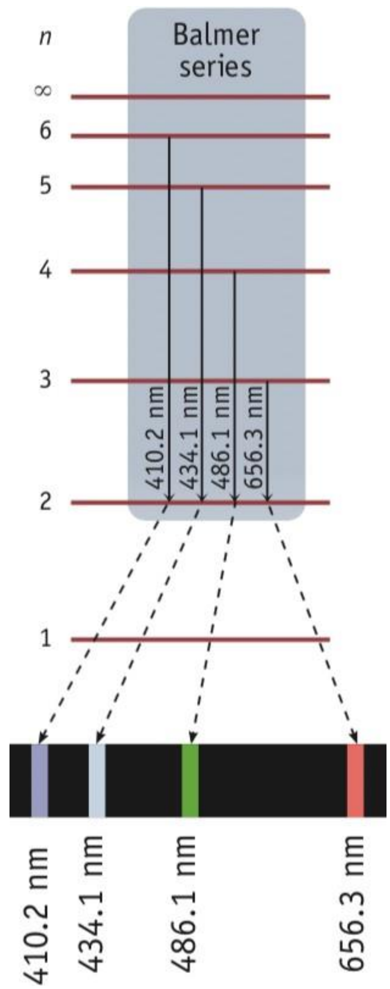

Figure 3. Diagram of the energy transitions in the hydrogen atom corresponding to the emission spectrum lines in the visible light region (between 380 nm to 780 nm). This emission series is called the Balmer series, first calculated by Johann Balmer in 1885. Note all transitions in this series have a value of n_f=2.

PREPARING FOR LAB 02: ATOMIC EMISSION SPECTRA

- This lab activity is composed of two parts. **Part I** focuses on measuring the emission spectra of different elements using a spectrometer provided in the lab. In **Part II**, you will use a simulation program on your computer to simulate the emission spectra you observed in Part I.
- **To prepare for Part I**: read through the lab procedure, and in your notebook create an outline of the procedure (don't copy verbatim!) – you should be able to perform the experiment from this outline. Also note down each measurement or calculation you must perform, leaving space to write down those numbers, equations, or data tables with proper units. Leave ample room to record observations at every step.
- Since you'll be observing the visible hydrogen emission lines during the lab (the Balmer series), it may be helpful to have the expected values for the wavelengths of those lines with their n_f and n_i values written somewhere in your lab notebook!
- **To prepare for Part II**: you do not need to have a procedure outline in your lab notebook for this part. Instead, download the instructions on your laptop and you will be working through the simulation as part of the post-lab worksheet. It is recommended that you still read through the procedure and prepare space in your notebook for measurements, data tables, calculations, and qualitative observations. Additionally, please install the following software on your computer ahead of time:

1. Install the (free) Wolfram Player on your computer. The link to install the player is: https://www.wolfram.com/player

   Some of the simulations used in the course are not compatible with the most recent versions (13.0.0 or newer) of the player. **Download version 12.3.1 (for PC) or 12.2.0 (for Mac) using the drop-down menu on the download page.**
2. Download the CDF simulations (if you haven't already for the lecture) that are posted at:
   http://quantum.bu.edu/CDF/101/2pyTo3dxyTransition.cdf
   http://quantum.bu.edu/CDF/101/1sTo2pTransition.cdf

EXPERIMENTAL, LAB 02: ATOMIC EMISSION SPECTRA

Safety

The atomic emission lamps in this lab work by discharging a high-voltage power supply into a glass tube containing the gaseous sample. Do not touch the metal connectors at either end of glass tubes while the power supply is turned on. Operating the lamps may cause them to heat up - do not touch the glass tubes while they are on or until they cool down after turning off.

The handheld spectrometers should not be pointed at any light source that could cause damage to an unprotected eye. Don’t use them to measure the spectrum of the sun or any light source that is painful to look at without the spectrometer.

Wearing gloves or washing your hands before and after is recommended when handling the spectrometers - they are not sanitized in between uses.

Materials

- Hydrogen emission lamp
- Helium emission lamp
- Other light source, such as fluorescent ceiling lights
- QA handheld spectrometer
- Mercury emission lamp

Procedure

Part I: Measuring the Atomic Emission Spectra

Use of the spectrometers: To use the handheld spectrometer, hold the narrow end of the instrument, the ocular window, close to your eye, and point the other end with the incident slit towards the light source to be measured, as shown in Figure 4. The diffraction grating in the spectrometer will split the light from source into its component wavelengths and project them on a scale within the instrument, as shown in Figure 5. The wavelength markings in the device are numbered every 100 nm with a single digit - for example, the line marked “4” in the readout corresponds to a wavelength of 400 nm. The line markings indicate every 10 nm, with a bold or different style of line on the 50 nm increment between numbered markings.

When reading the wavelength of a spectral line from the handheld spectrometer, match up the approximate center of the line to a position directly above it on the wavelength scale, and estimate that position as best as you can between the markings on the wavelength scale. *For group discussion:* Based on the markings shown in Figure 5, how may significant figures should you record for the wavelength in nm?

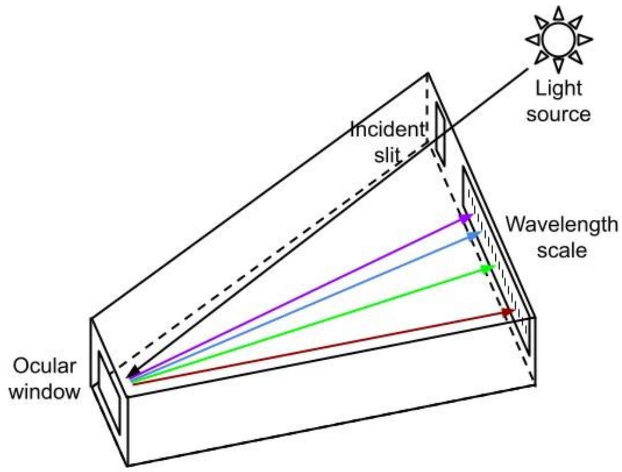

Figure 4. Diagram of a QA handheld spectrometer. For proper use, the ocular window should be held near the eye and the incident slit should be pointed at the light source. The spectrum can then be read out under the wavelength scale.

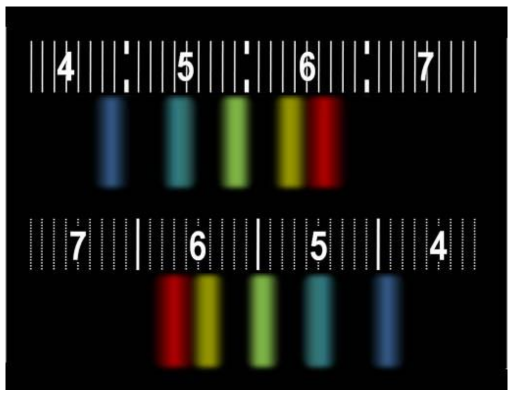

Figure 5. Example spectral readouts from a QA handheld spectrometer. The numbers correspond to hundreds of nanometers, and the lines mark out increments of 10 nm.

Measuring the spectra:

1. Have everyone in the group use the spectrometer to look at a light source other than the atomic emission lamps, such as the overhead lights in the lab. Qualitatively describe the spectrum of the light source you observed in your lab notebook.
   For group discussion: how and why is the spectrum different from what you expect the atomic emission spectra to look like?

2. Turn on one of the **hydrogen** emission lamps, and look at it through the spectrometer. Adjust the distances between your eye, the spectrometer, and the lamp until you can clearly see multiple emission lines on the spectrometer. It may help to shield the incident slit on the spectrometer from the overhead lights or other light sources in the room with your hand that is not holding the spectrometer.

   Note: If another group is using the hydrogen lamp, try a different lamp and come back to the hydrogen lamp later.

3. Make sure you aren't observing spectral lines from any other sources - you may ask your instructor to turn off the overhead lights, or you can have a group member turn the lamp

off and back on again so you can tell which lines are from the lamp. Record the color and wavelength of the atomic emission lines that you can observe using a data table. Make sure you note down which element you were observing in the header of the table!

Write down any other qualitative observations about the spectrum you might have, for example, were any emission lines much brighter or dimmer than the others? Were some wideror thinner?

4. Have everyone in your group repeat step 2 and 3 and discuss the spectral lines that you observed and their wavelengths. Match up as many lines as you can between your measured spectrum and those of your group members and write down what wavelengths your group members estimated for the same spectral lines in your data table. Calculate the average wavelength for each line and note it down on your data table to use for the data analysis later.
5. Repeat steps 2 to 4 for another atomic lamp that is not hydrogen.

At the end of this Part I, you should have recorded the emission wavelengths for the hydrogen lamp, and one other lamp of your choice.

Turn off all the atomic emission lamps (as long as someone else isn't using them) and return the spectrometer to your instructor.

Obtain a copy of the post-lab assignment from your instructor to work through the following simulations on your computer and answering questions in the post-lab worksheet.

Part II: Simulating the Atomic Emission Lines

Simulating the Balmer series:

You do not need to write these instructions in your lab notebook but you should save them to the computer on which you're running the simulations.

1. Open up the Wolfram player on your computer and load the first CDF simulation (http://quantum.bu.edu/CDF/101/2pyTo3dxyTransition.cdf). Click "Enable Dynamics" in the player so that the simulation will run on your device.
2. Start by tuning the light frequency all the way to the left so that v/v_0 is set to 0.25. Once the frequency has been set, press the play button (▶).
3. Discuss with your group the following questions and answer them in the post-lab worksheet:
   a) What do the arrows moving back-and-forth (left-to-right) represent in the simulation?
   b) What does the red dot in the middle of the simulation area represent?
   c) The blue dumbbell shape represents a hydrogen atom *2p* electron wave. What happens to this wave when the simulation is running at a light frequency that is too low (v/v_0 = 0.25)? Explain your observation in no more than one sentence.
4. Set the Transition Progress slider (top slider) all the way to the right, then click the arrow pointing to the right (→) to reverse the direction of the simulation (the arrow should now be

pointing to the left ←). Then re-tune the light frequency so that the simulation is run at the resonant frequency of the electron wave (i.e., $v = v_0$ or $v/v_0 = 1.0$).

5. Discuss with your group the following questions and answer them in the post-lab worksheet:
   a) At the end of the simulation, the electron wave changed from a 3*d* electron wave to a 2*p* wave. Did the electron gain or lose energy during the process? What about light?
   b) Is this process instantaneous or gradual?
   c) The energy change of electron wave during the process can be modeled using the following equation: $\Delta E_{electron} = E_f - E_i$. What is the value of n_f and n_i for the process you observed here?
   d) Based on your answer in part c), calculate $\Delta E_{electron}$. Does this answer match your conclusion in part a)? For hydrogen atom electron cloud: $E_n = (-2.1799\ \mathrm{aJ})\left(\frac{1}{n^2}\right)$.
   e) During the process, is light being absorbed or emitted? What is the wavelength and color of this light?
   f) Did you observe this line in your spectrometer measurement of the hydrogen lamp in **Part I**? If so, which line is it?

Simulating the Lyman series:

1. Load the second CDF simulation into the Wolfram player. (http://quantum.bu.edu/CDF/101/1sTo2pTransition.cdf)
2. Set up the simulation to run in **reverse** with the light frequency set to be resonant with the electron wave (similar to step 4 in the simulation of the Balmer series line), and run it.
3. The initial and final states of the electron wave in this simulation are 2*p* and 1*s* waves of a hydrogen atom. Discuss with your group and write down your answers to the following on the post-lab worksheet:

a) Draw an energy diagram that represents the process you observed.

b) Calculate the wavelength of the light emitted during the process. Will you be able to see it?