Experiment 40B

LIGHT, ENERGY AND SPECTRA

MATERIALS: Hydrogen vapor lamp and power supply; Neon vapor lamp and power supply; Helium vapor lamp and power supply; quantitative spectroscopes (4); LED circuit board (6); adjustable voltage DC power supply (6); incandescent light bulb fixture (1)

PURPOSE: The purposes of this experiment are: (1) to observe the relationship between wavelength, frequency and energy of light; (2) to observe the emission spectrum of atomic hydrogen and (3) to determine the wavelengths and energies of some of the electronic transitions of the Balmer series for hydrogen.

LEARNING OBJECTIVES: By the end of this experiment, the student should be able to demonstrate these proficiencies:

1. Calculate the energy of light from its wavelength.
2. Calculate the energy of light from its frequency.
3. Calculate the wavelengths of light expected for specific electronic transitions of hydrogen.
4. Compare the expected wavelengths of the hydrogen atom to the observed wavelengths.
5. Determine the value of Planck's constant.

DISCUSSION:

Much of what we know about the atomic or molecular nature of the world has been obtained by the study of how radiation interacts with matter. When people hear the term "radiation" they commonly think of *ionizing* radiation associated with nuclear processes. We shall study that in SC112, but now we will focus on the radiation we all use and are exposed to daily; mainly the visible part of the electromagnetic spectrum we call light. Light is a wave phenomenon, characterized by frequency, ν, and wavelength, λ, which are related to each other through the speed of light, c.

$$c \ = \ \lambda\,\nu \ = \ 2.998 \text{ x } 10^8 \text{ m/s} \tag{1}$$

Planck showed that the energy of light is related to its frequency,

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

where the proportionality constant, $h = 6.626 \text{ x } 10^{-34}$ J s, was named in his honor.

In our everyday activities, we perceive different wavelengths of visible light as an assortment of colors. Any specific wavelength will correspond to only one of these colors, although as a practical matter our eyes actually are responding to broad ranges of wavelengths. Additionally, many colors we see actually result from familiar additive or subtractive combinations. "Additive color synthesis" is the creation of color by mixing colors of *light*. For example, flat-panel LED (light-emitting diode) computer screens or TVs mix red light from red-emitting LEDs with green light from green-emitting LEDs to make yellow light on the screen. There are no yellow LEDs in the device. By contrast, the color we see on painted walls or the printed page results from "subtractive color synthesis", where color is created by mixing colors of *pigment*. The pigments absorb (remove) some colors of white light, and we see what was transmitted or reflected (i.e., what was *not* absorbed). What we see on dollar bills is essentially white light seen through yellow and blue dyes – the green color is what reaches your eye after the blue and yellow dyes remove some wavelengths of the full white light spectrum. (See https://cs.nyu.edu/courses/fall02/V22.0380-001/color_theory.htm for more details.) Colors absorbed or transmitted are complementary, appearing on opposite sides of a color wheel. Our perceptions of colors are also influenced by the relative intensities of the light, the relative absorptivities of the various wavelengths by pigments or dyes, and the responses of detectors (including our eyes), but the basic ideas of emission, transmission and absorption of light are parts of everyday life.

We have seen that the frequency of light is associated with the energy of the light, and indeed Einstein treated light as a

collection of particle-like packets of energy now known as "photons". When an atom or molecule absorbs or emits light, it gains or loses the same amount of energy as is carried by the photon. The absorption of light by an atom or molecule always represents that molecule going from a lower energy state (before it accepts the photon) to a higher energy state (after the photon is absorbed). Similarly, when an atom or molecule emits light, it goes from a higher energy state to a lower one (Fig. 1).
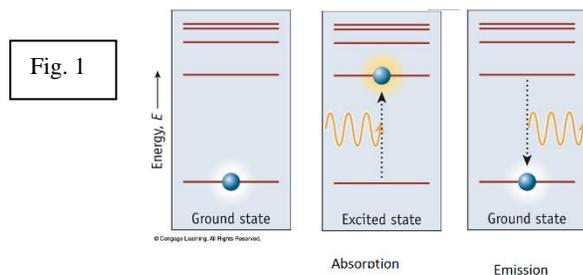
When atoms or molecules change their energy in response to absorption or emission of a photon, they undergo a "transition". Only one such transition is shown in Fig. 1, but many transitions are possible for most substances, which means that many different photons of different energies (and frequencies) could be involved. A "spectrum" is a plot showing the collection of transitions an atom or molecule can undergo. Atomic and molecular spectra are characteristic of the substance, so we shall have many opportunities to use them this year. But they basically fall into a few categories (Fig. 2):

Fig. 2

Category	Type of spectrum	Example	Comments
Emission	Continuous		All wavelengths emitted. Typical of incandescent objects (heated until they glow)
Emission	Line or band		Only a few specific wavelengths emitted. Typical of isolated atomic or molecular species.
Absorption or Transmission	Line or band		Most wavelengths transmitted; only a few removed (absorbed). Typical of isolated atomic or molecular species exposed to a continuous spectrum of light.

(color spectrum images from http://tap.iop.org/atoms/quantum/501/page_47004.html, Institute of Physics; accessed 6/27/2016)

For any given substance the absorption spectrum will involve the same wavelengths of light being absorbed as are emitted in the emission spectrum of that substance, because the same energy levels are connected by the transition (see Fig. 1). Transmission spectra will show wavelengths *not* absorbed.

Light emission from LEDs also occurs when the system goes from high to low energy, but rather than discreet (specific) energy levels, bands of energy levels of atoms in the crystal are involved. The bands are separated by an energy known as the "band gap". When sufficient voltage is applied to move electrons into the higher energy band, they can drop down to lower energy "holes", releasing light as they do so. The voltage required to accomplish this depends on the size of the band gap, so the minimum voltage to light the LED corresponds to the color observed.

Emission Spectra of Elements

When an electron in an atom undergoes a transition from a higher energy level to a lower energy level, the atom emits light at a discrete frequency and wavelength determined by the energy difference between the levels Because the separation between the energy levels depends on the type of atom, the emission spectrum is characteristic of the element. Hydrogen atoms, if excited by an electrical discharge, emit a series of lines in the visible region called the Balmer series. This series corresponds to transitions from several different excited states to the n = 2 level. Three lines of the Balmer series can be observed with the unaided eye.

A diffraction grating is a transparent film ruled with a number of closely spaced grooves. It is used to separate the light from the emission lamp according to its wavelengths. If light from an incandescent lamp is directed onto the grating, a continuous spectrum of colors is formed. The grating produces an image of the light for each color emitted; because all colors are emitted, these images blur together and appear as a continuous band (like a rainbow). When a hydrogen lamp is viewed through the grating, only three images of the light will appear, each in a different color. These correspond to the individual emission lines of the Balmer series, each with a different wavelength. Because they are separated in space, the images appear distinct. The first two images in Figure 2 were produced using such a grating.

As indicated above, the emission spectra of hydrogen atoms and "hydrogen-like" (one electron) ions consist of a set of individual "lines" of specific wavelengths. The numerical values of the wavelengths of these lines fit a particular mathematical pattern called a series. Rydberg determined that the equation

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

(3)

could reproduce the pattern observed by Balmer for hydrogen, where R_H is the Rydberg constant, $R_H = 1.0974 \times 10^7$ m^{-1}, and n are positive integers for the states in the emission process. In the Balmer series n > 2. While Rydberg did not understand what the numbers meant, this high regularity and agreement with experiment convinced him that he was on the right track. The wavelengths that you observe in this experiment should satisfy the Balmer-Rydberg relation within experimental error. Knowing the relationship, you can calculate wavelengths for a variety of integer pairs and compare those with your experimental data. A close match between observed and calculated values will allow you to assign the quantum numbers for each transition.

PROCEDURE:

Note: Data collection for the three parts of this lab can be done in any order. Your instructor may assign a specific order or rotation to minimize delays in some parts. Follow the directions of your instructor. When not collecting data, you should work on analysis of the data you already have.

Part A. Color, Wavelength and Energy

1. At your lab station locate the DC power supply and the LED circuit board. (Handle the circuit board *gently*!) On the power supply, rotate both the voltage and current knobs to the left (counterclockwise), and then turn on the power supply. Both readouts should indicate zero.

2. Connect the black (negative) wire from the power supply to the <u>single</u> post on the circuit board set off from the row of LEDs. Connect the red (positive) wire from the power supply to the post just below the right–most LED. See photos at right → .

3. On the power supply, the indicator light marked "CC" next to the lower (current) knob should be glowing red. Rotate the current (lower) knob just slightly to the right, until the CC indicator light just turns green. <u>Leave the current knob at this position for the remainder of the experiment.</u>

4. Rotate the upper (voltage) knob *slowly and slightly* to the right, and continue until the right-most LED just begins to glow. Leave off the lights in your hood and look at the LED <u>from the top</u> for best sensitivity. Now, slowly rotate the voltage knob to the left until the LED just goes out. Repeat this step to refine the measurement until you have reached the lowest possible voltage at which the LED barely glows. Record the LED voltage from the upper readout on the appropriate color line in the Data Table.

5. Rotate the voltage knob back to zero volts. Disconnect the red wire from the LED circuit board and move it to the contact post one LED to the left. You can leave the black wire connected.

6. Repeat steps 4 and 5 until you have measured the minimum voltages required to excite each of the six LEDs.

7. Turn both knobs fully counterclockwise, disconnect the wires from the circuit board, and turn off the power supply.
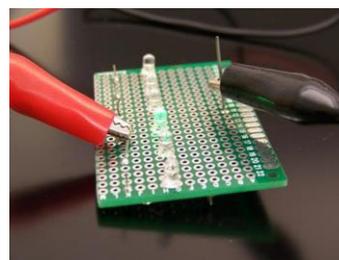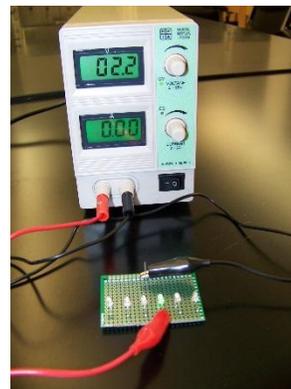
Part B. Emission Spectra

1. Turn on the incandescent light bulb. Use the hand-held spectroscopes to examine the emission spectrum of the light bulb. To use the spectroscope, hold it level to the floor and peer in the opening at the narrow end while pointing the slit at the left side of the wide end towards the light source. The spectrum will be apparent on the screen to the right. The wavelength scale visible on the screen is calibrated in hundreds of nanometers (i.e.; "6" means 600 nm, etc.)

2. Now aim the spectroscope at the overhead (fluorescent) room lights. The spectrum should appear similar, but also have additional features. Record your observations on the Data page, describing the similarities and differences between the spectra of the two light sources, and estimating the wavelengths (in nm) of the purple line and the bright green line in the spectrum of the overhead lights. These lines are due to the mercury vapor in the fluorescent bulbs.

Part C. Quantitative Determination of the Hydrogen Atom Emission Spectrum

CAUTION: Be sure to wear goggles as protection from ultraviolet radiation given off by the emission tube.

1. Turn off the overhead room lights; light from adjacent rooms should be sufficient. Now aim the spectroscope at the hydrogen atom lamp. You should see three lines, one red, one blue-green and one violet. Estimate the wavelengths (in nm) of these lines. Record your values in the Data section.

2. Have your lab partner independently estimate the wavelengths of these lines. Record your partner's values in the Data section.

3. Turn off the hydrogen lamp and turn on the helium lamp. Observe the spectrum through the spectroscope. Record your observations on the Data page, describing the similarities and differences between the spectra and estimating the wavelengths (in nm) of the bright features in the spectrum of the helium lamp.

4. Turn off the helium lamp and turn on the neon lamp. Observe the spectrum through the spectroscope. Record your observations on the Data page, describing the similarities and differences between the spectra of the two light sources, and estimating the wavelengths (in nm) of the bright features in the spectrum of the neon lamp.

Name _____   Section _____

Partner _____   Date _____

DATA SECTION
Experiment 40B

Part A. Color, Wavelength and Energy

Color	Wavelength (nm)	Minimum voltage to excite LED (V)
violet	397.5	
blue	452.5	
green	520.0	
yellow	590.0	
orange	606.0	
red	622.5	

Part B. Emission Spectra

1. Observations on spectra of incandescent and fluorescent light sources (similarities, differences, etc.)

_____

_____

Estimated wavelengths of bright features in fluorescent light spectrum:

Purple line: _____

Green line: _____

Part C. Quantitative Determination of the Hydrogen Atom Emission Spectrum

1. Observations on spectra of hydrogen atom lamp (similarities, differences, etc.)

_____

_____

_____

Estimated wavelengths of bright features   First reading   Second reading

red   _____   _____

blue-green   _____   _____

violet   _____   _____

2. Observations on spectra of helium atom lamp (similarities, differences, etc.)

_____

_____

_____

Colors and estimated wavelengths of bright features

_____

_____

3. Observations on spectra of neon atom lamp (similarities, differences, etc.)

_____

_____

_____

Colors and estimated wavelengths of bright features

_____

_____

DATA TREATMENT
Experiment 40B

Part A. Color, Wavelength and Energy

(A.1) Create a spreadsheet with the following columns: Color; Wavelength (nm); Voltage (V); Frequency (s^{-1}) and Energy (J). Enter your color, wavelength and minimum excitation voltage data.

(A.2) Use spreadsheet formulas (NOT a calculator) to calculate the frequency of the light emitted by the LED from the wavelength of the light (Eq.1). SHOW your work for the red diode.

(A.3) Use spreadsheet formulas (NOT a calculator) to determine the energy corresponding to the applied voltage (Eq. 2). The electrical work required to move the electrons between the bands is the product of the voltage and the charge on an electron: energy (J) = voltage (V) * 1.602×10^{-19} J/V. SHOW your work for the red diode.

(A.4) Make two separate plots: *Energy vs. Wavelength* and *Energy vs. Frequency*. Include a linear trendline on each plot.

        Which plot is *best* described by a linear trendline? (circle one)   *Energy vs. Wavelength*   *Energy vs. Frequency*

        What are the values of slope and intercept for that plot? Be sure to include units for each.

                slope: _____   intercept: _____

(A.5) Note by comparison with Eq. 2, the slope of the line should be an experimental estimate of the value of Planck's constant. Calculate the percent error from the accepted value of 6.63×10^{-34} J s.

Part B and C. Quantitative Determination of the Hydrogen Atom Emission Spectrum

(C.1) Enter your wavelength values for the fluorescent lights into an Excel spreadsheet. The purple line is known to have a wavelength of 436 nm and the green line is known to be at 546 nm. Use these values to calibrate your spectroscope. Determine how far each line is from the known value. Deviation from known values: purple_____ nm and green _____ nm.

(C.2) Enter your wavelength values for the hydrogen lamp lines into the spreadsheet. Apply the wavelength correction you calculated in step C.1 to these values.

(C.3) Using the Rydberg equation for the Balmer series, determine the values of n for each of these wavelengths. These should be integers so round your results to the nearest whole number.

(C.4) It is known that these lines correspond to the values of n = 3, n = 4, n = 5. Use Excel to calculate the expected wavelengths for these quantum numbers and determine the percent error for your results.

1. Which emission line in the Balmer series has the *longest* wavelength? _____

Which emission line in the Balmer series has the *highest* energy photons? _____

What type of mathematical relationship exists between energy and wavelength? _____

2. As quantum mechanics developed, it eventually became clear that the form of the Balmer-Rydberg equation corresponded to a difference in the energies of the two stable states connected by the electronic transition, i.e.

$$\frac{1}{\lambda} = R\left(\frac{1}{n_{final}^2} - \frac{1}{n_{initial}^2}\right)$$

Since $\Delta E = E_{final} - E_{initial}$, this becomes:

$$\Delta E = -Rhc\left(\frac{1}{n_{final}^2} - \frac{1}{n_{initial}^2}\right)$$

where $\otimes E$ is the energy of the photon emitted in the transition, h is Planck's constant, c is the speed of light and n_{final} and $n_{initial}$ are the states involved in the transition. Calculate the value (in Joules) for a transition from $n_{initial} = 6$ to $n_{final} = 2$ for a Balmer series transition. Calculate the wavelength of light emitted by this transition. Is this wavelength in the visible portion of the electromagnetic spectrum (~400-700 nm)?

Energy: _____ Wavelength _____ Visible? (yes or no)_____

3. Both He$^+$ and Li^{2+} are "hydrogen-like" ions, in that they only have one electron. These ions will also produce a line spectrum that obeys the Balmer-Rydberg equation, but with different R constants (we will call them R_{He} and R_{Li}). Knowing the wavelengths, and the appropriate integers for $n_{initial}$ and n_{final}, you can calculate these constants and gain some additional physical insight.

       a) In the He$^+$ spectrum, a line appearing at 164.1 nm corresponds to the red emission you observed for H (i.e., the 164.1 nm line for He$^+$ has the <u>same</u> values of $n_{initial}$ and n_{final} as does the red line of H). Use that information to calculate the constant R_{He} for the helium ion. Show your work. Record that value in the table below.

       b) Repeat the calculation for the Li^{2+} ion spectrum, where a line appearing at 72.9 nm corresponds to the red emission you observed for H (i.e., the 72.9 nm line for Li^{2+} has the same values for $n_{initial}$ and n_{final} as does the red line of H). Record the value of R_{Li} in the table below.

Constant	R_H (for Hydrogen Atom)	R_{He} (Helium Ion)	R_{Li} (Lithium Ion)
R value	1.0974 x 10^7 m^{-1}		
Integer	1		

4. You should find that the constants R_{He} and R_{Li} are integer multiples of the Rydberg constant R. Show the values of these integer multiples in the table above. The integers for all three are related to the atomic structure of the specific atoms (H, He, or Li). How do the integers relate to the <u>atomic structures</u> of the atoms H, He, and Li? (HINT: focus on the nucleus.)

Name _____   Section _____

Date _____

PRE-LAB QUESTIONS
Experiment 40B

1a. A green laser pointer emits light with a wavelength of 532 nm. What is the frequency of this radiation, in Hz?

1b. What is the energy, in joules, of the 532 nm photon emitted by a green laser pointer?

1c. A watt (W) is a unit of power equal to 1 joule per second. How many 532 nm photons would be required to hit a surface in one second to deliver 100.W; the power rating of a typical light bulb?

2. List the following regions of electromagnetic radiation according the increasing energy: infrared, microwave, radio, visible, X-ray.

lowest Energy _____   _____   _____   _____   _____ highest Energy