

Experiment 40

9/28/18

LIGHT, ENERGY AND SPECTRA

MATERIALS: hydrogen lamp; diffraction grating (transmission type); meter stick; optical rails (2); quantitative spectroscopes (4); LED circuit board; adjustable voltage DC power supply; demo AS-13 flame test kit; Spectronic 200 spectrometer; 100 mL graduated cylinders (2); concentrated food colors in dropper bottles (red, yellow, green, blue); incandescent light bulb fixture

PURPOSE: The purposes of this experiment are: (1) to observe absorption, emission and transmission spectra of a variety of sources; (2) to evaluate the energy-frequency relation to determine Planck's constant; (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. Understand the concepts associated with absorption, emission, and transmission spectra.
2. Understand the relationships between energy, frequency and wavelength of light.
3. Calculate the wavelengths of light expected for specific electronic transitions of hydrogen.
4. Create a simple Excel spreadsheet for the execution of repetitive calculations.

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. The "radiation" discussed here is not the *ionizing* radiation associated with nuclear processes, but rather the radiation we all use and are exposed to daily, 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 .

$$c = \lambda \nu = 2.998 \times 10^8 \text{ m/s} \quad (1)$$

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

$$E = h \nu \quad (2)$$

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

Objects may release ("emit") their own light, or remove ("absorb") or reflect light directed at them from an external source. 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 our eyes actually are responding to broad ranges of wavelengths. Additionally, many colors we see 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 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 passed 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. A molecule that absorbs red light will transmit (and appear) green. 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.

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 (Figure 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 (Figure 2).

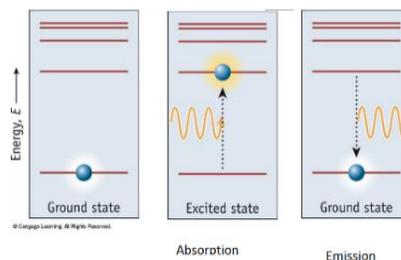


Figure 1. Comparison of absorption and emission

Figure 2. Categories of spectra

Category	Type of spectrum	Example	Comments
Emission	Continuous		All wavelengths emitted. Typical of incandescent objects (heated until they glow)
Emission	Line		Only a few specific wavelengths emitted. Typical of isolated excited atomic or molecular species.
Emission	Band		Many broad sets of wavelengths emitted, as shown by the peaks in the chart. This is the emission spectrum of a white “glowstick”
Absorption	Line or band		Most wavelengths transmitted; only a few removed (absorbed) –where dark lines or peaks appear. Typical of atomic or molecular species exposed to a continuous spectrum of light. Note that wavelengths absorbed will be the SAME as wavelengths emitted for a substance
Transmission	Band		Note that the transmission spectrum is like the inverse of the absorbance spectrum of the substance.

Color atomic spectrum images from http://tap.iop.org/atoms/quantum/501/page_47004.html, Institute of Physics; accessed 6/27/2016; glowstick spectrum from https://en.wikipedia.org/wiki/Glow_stick#/media/File:White_glow_stick.png, Wikipedia, accessed 6/19/2017.

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 Figure 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 discrete (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.

White light is a collection of all visible colors. In order to actually separate the colors from each other, we need a device that disperses white light into its components. Traditionally a prism was used, but today a “diffraction grating” is more common. This is a transparent (or reflective) material ruled with a number of very closely spaced lines. As a result of wave interference effects at the rulings, white light passing through the grating is separated (in space) into its constituent wavelengths according to the expression

$$\lambda = d \sin \theta \quad (3)$$

where the angle θ is dictated by the geometry of the setup and d is the separation between rulings on the grating. Using such a grating to examine the emission spectra of simple 1-electron atomic/ionic species, the wavelengths of light absorbed or emitted can be described by a simple expression that relates to the quantum numbers of the energy levels involved in the transition.

$$\frac{1}{\lambda} = (R_H) \left(\frac{1}{m^2} - \frac{1}{n^2} \right) \quad (4)$$

where m and n are the quantum numbers of the lower and upper state, respectively and R_H is the Rydberg constant, which for H atoms is $1.0974 \times 10^7 \text{ m}^{-1}$. We will use this expression to analyze the line emission spectrum of the hydrogen atom.

PROCEDURE:

Note: Data collection for the four 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 approximately zero.
2. Connect the black (negative) wire from the power supply to the single 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 Figures 3 and 4.
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 red (CC) light goes out and the green (CV) indicator light comes on. *Leave the current (lower) knob at this position for the remainder of the experiment.*
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 from the top for best sensitivity.) Now, slowly rotate the knob to the left until the LED just goes out. Repeat this step to refine the measurement until you definitely have 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.

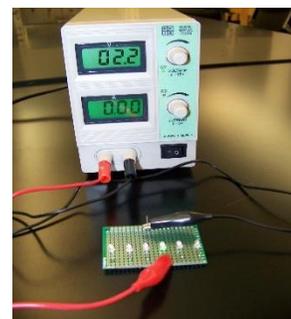


Figure 3. DC power supply and LED circuit board

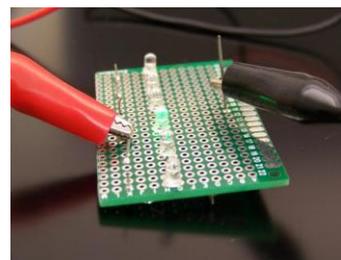


Figure 4. LED circuit board

Part B. Color, Absorption Spectra and Transmission Spectra

1. Carefully rinse out, and then fill each of the two graduated cylinders at your station to the 100.0 mL mark with deionized water. (Use a plastic dropper to get the meniscus exactly to the line.) Use these to prepare two solutions of food coloring dye as follows:
 - odd numbered hoods – prepare red and green solutions – add ONE drop of the dye to one graduated cylinder; repeat with the second dye and second graduated cylinder.
 - even numbered hoods – prepare yellow and blue solutions– add ONE drop of the dye to one graduated cylinder; repeat with the second dye and second graduated cylinder.Use your glass stirring rod to stir each solution; wipe off the stirring rods between solutions to avoid cross-contamination.
2. Turn on the Spectronic 200 spectrometer and proceed to the SCAN menu. Set the instrument to scan the visible spectrum, from 400 nm to 700 nm in *Absorption* (ABS) mode. Rinse a plastic cuvette with deionized water; then rinse with the test solution, and finally fill it about 2/3 full with one of your solutions. Insert the cuvette in the spectrometer and scan the spectrum. Use the cursor keys to find the wavelength(s) of maximum absorption (where the plot peaks). Record that wavelength and make a rough sketch of the absorbance spectrum in the Data Table. Repeat with your second dye solution.

- Return to the SCAN menu and set the instrument to record in the *Transmission (%T)* mode. Scan each of your solutions and sketch the Transmission spectra in the Data Table. Note the inverse appearance.
- From a nearby hood, obtain samples of the other two dye solutions, and repeat steps 2 and 3 with those.
- Go to the Cengage website

http://www.cengage.com/chemistry/discipline_content/dvd/Power_Lectures/General_Chemistry/dswmedia/Presentation_Simulations/plancks_equation_s.swf

Use your four solutions (in cuvettes or graduated cylinders) and estimate the apparent visual wavelength by comparison with the online spectrum. (Just click on the screen spectrum at the color you think provides the best match.) Read the wavelength listed and record in the Data Table. (Think – is the color you see the one(s) absorbed, or the one(s) transmitted? Should this match the maximum absorbance, or maximum transmittance?)

- When finished, empty your graduated cylinders and cells in the sink, and rinse each with deionized water.

Part C. Emission Spectra

- In the small fume hood, 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. (See Figure 5, which is a top view.) The spectrum will appear 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.)

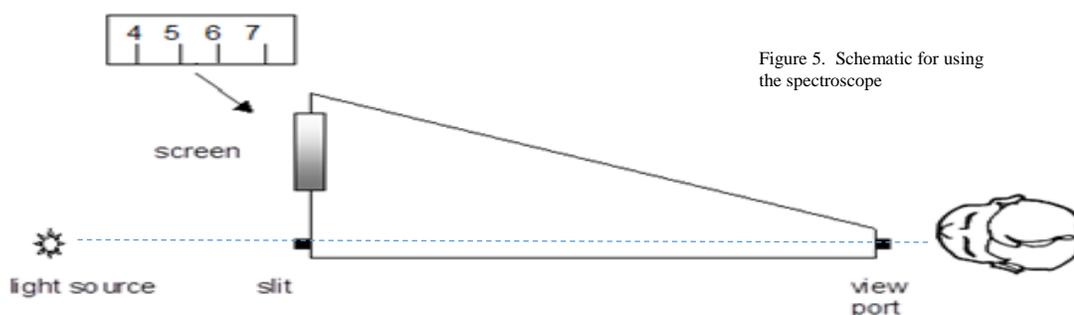
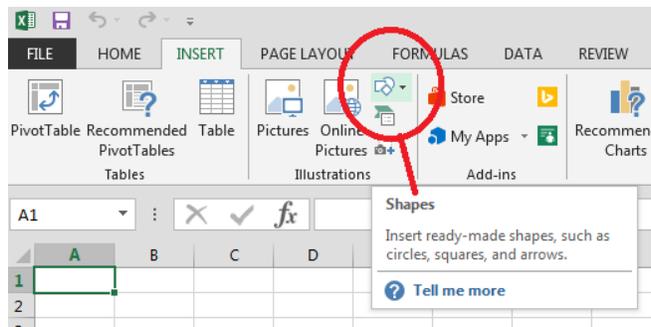


Figure 5. Schematic for using the spectroscope

- 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 bright features in the spectrum of the overhead lights.
- Open an Excel spreadsheet and use **Insert/Shapes** to create a red box, a green box, a yellow box and a white box (each filled in with that color). Make each shape occupy about one fourth of the screen.
- Use the hand-held spectroscope to examine the emission spectra of these boxes. Record your observations. You can delete the shapes after viewing them.
- Instructor Demonstration** – Your instructor will demonstrate a flame test for several salts. Observe and record the color of the flame which is associated with the emission from the metal ions in the salt. Note that most flames will have some yellow color, but **DO NOT** report that except for the sodium salt. Look for the more subtle (and fleeting) colors.



6. Write down a typical wavelength associated with that color. (You will NOT be able to measure them with the spectrosopes because the colors are too dim and short-lived. Just give typical values from your text or the Cengage website

http://www.cengage.com/chemistry/discipline_content/dvd/Power_Lectures/General_Chemistry/dswmedia/Presentation_Simulations/plancks_equation_s.swf

Part D. 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.

The apparatus will already be assembled. A diagram appears below in Figure 6.

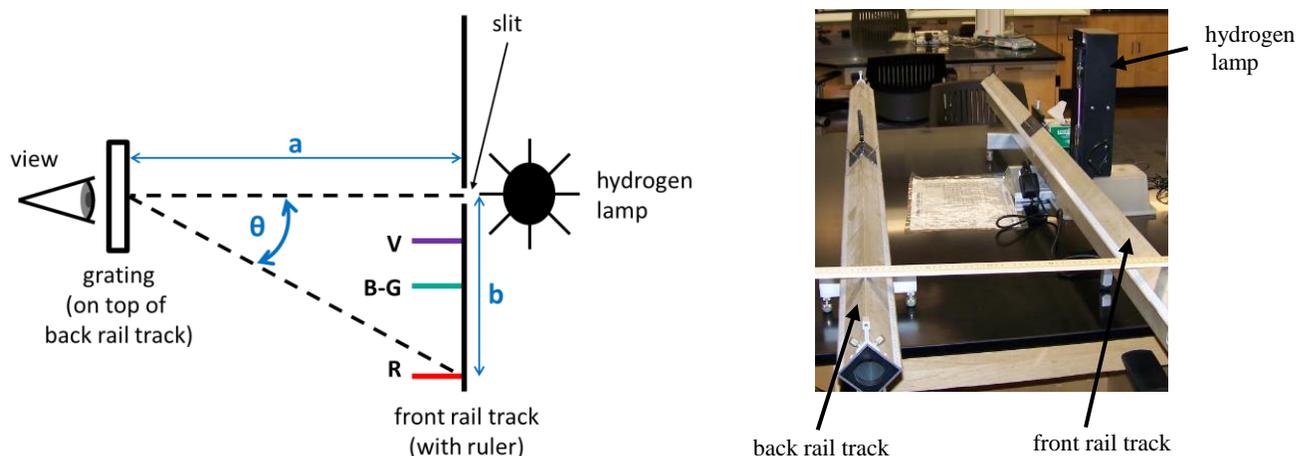


Figure 6. Schematic and picture of the hydrogen lamp apparatus. V, B-G and R stand for violet, blue-green and red; they are the diffracted line images of the lamp. Note that the angle θ depends on the position of the line image. The distances a and b are described in the text and are used for calculating the angle θ

Record all readings, to the proper number of significant figures and with appropriate units, in the Data Section.

1. Make sure that the lamp is centered at 100 cm on the front rail track. Check that the diffraction grating is centered at 100 cm on the back rail track. Check that the two rail tracks are 50 cm apart at both ends.
2. Turn off the overhead room lights above setup; light from adjacent fixtures should be sufficient. Look at the images on the diffraction grating; three colored images should be present to one side of the straight-through view of the lamp (Figure 7). Using your line of sight, align the red line emission with the marker at the center of the diffraction grating. Have your partner move the needle indicator on the front rail track to be in line with the emission image, as viewed through the grating. Record the position of the red line/needle indicator on the front rail scale to the nearest 0.1 cm.
3. Repeat for the Blue-Green and Violet lines.

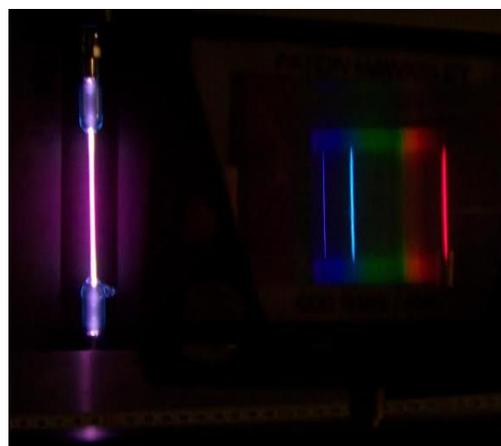


Figure 7. Picture of emission from the hydrogen lamp (bright image at left) and three diffracted emission lines.

Name _____

Section _____

Partner _____

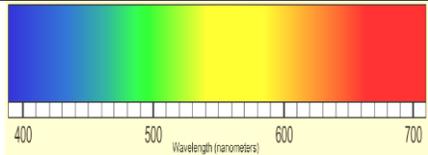
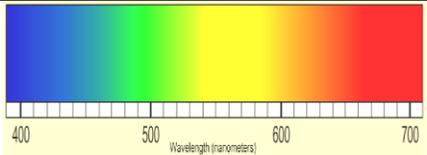
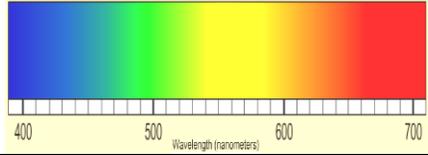
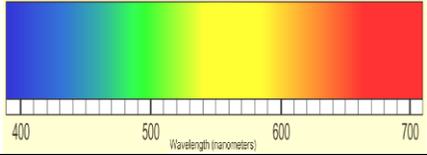
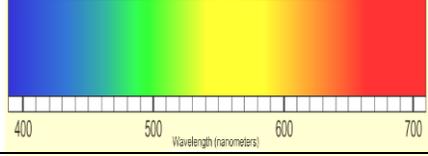
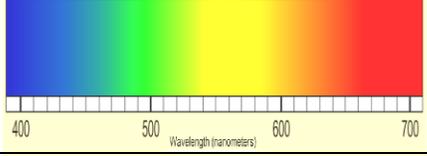
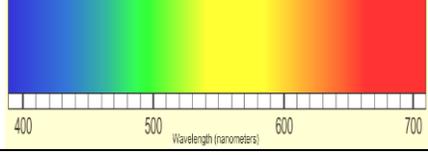
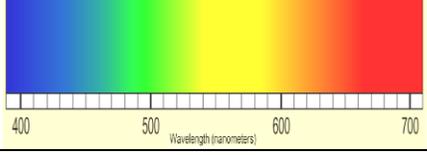
Date _____

**DATA SECTION
Experiment 40**

Part A. Color, Wavelength and Energy

Color	Wavelength (nm)	Minimum voltage to excite LED (V)
violet	398	
blue	456	
green	513	
yellow	596	
orange	613	
red	640	

Part B. Color, Absorption Spectra and Transmission Spectra

Dye Color	Maximum Absorbance Wavelength (nm)	Absorption Spectrum Sketch (Draw on the spectrum)	Transmission Spectrum Sketch (Draw on the spectrum)	Apparent Visual Wavelength (nm)
				
				
				
				

Part C. Emission Spectra

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

Estimated wavelengths (in nm) of bright features in fluorescent light spectrum. (Note that the scale reads in hundreds of nm, i.e. "6" = 600 nm):

2. Observation of computer displays:

red box	green box
yellow box	white box

3. Flame test observed emission colors:

Salt	Color	Typical Wavelength (nm)

Part D. Quantitative Determination of the Hydrogen Atom Emission Spectrum

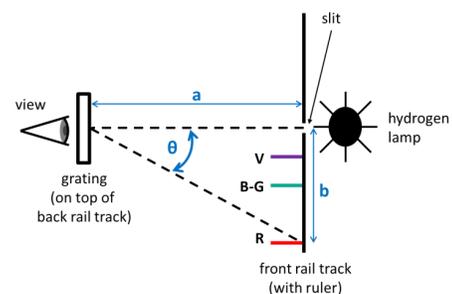
Diffraction Grating Ruling _____ 600 _____ lines / mm

Number of separations/mm in diffraction grating _____

Calculate width of one separation, **d**, in nm _____

Lamp position along front rail track _____ 100.0 cm _____

Distance (a) from front rail track to grating _____ 50.0 cm _____



Colored images seen through grating	color	position on front rail (cm)	distance (b) from lamp position (cm)
	red (R)		
	blue-green (B-G)		
	violet(V)		

DATA TREATMENT
Experiment 40

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 determine 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 in Joules corresponding to the applied voltage. 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*. (Energy is on the y-axis in each.) Include a linear trendline on each plot, and the equation and R^2 value. Note: If the equation is not displayed with at least three significant figures in the slope, format the trendline equation label so that three significant figures are displayed.

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

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

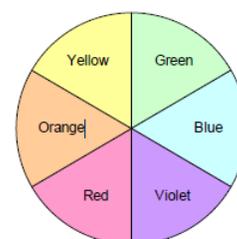
slope: _____ intercept: _____

(A.5) Note by comparison with Equation 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. Color, Absorption Spectra and Transmission Spectra

(B.1) Qualitatively describe how the Absorption spectrum and the Transmission spectrum of the same substance compare to each other. (Do NOT give an equation as your answer!)

(B.2) Qualitatively describe how the observed color of a substance compares to the color absorbed by that substance. You might find a "color wheel" useful in your answer.



Part C. Emission Spectra

(C.1) You should have observed a green color in the flame emission spectrum of copper salts, and a yellow color in the flame emission spectrum of sodium salts. Typical strong *emission* features from these species are listed below.

Metal	Emission wavelength (nm)	Absorption wavelength (nm)
Cu	520	
Na	589	

Predict the wavelength of light that you would expect to be *absorbed* (i.e. NOT transmitted – see Figure 2) by each of these species; enter in the table above. Explain your answer.

Part D. Quantitative Determination of the Hydrogen Atom Emission Spectrum

(D.1) You need to know the separation **d** between lines on the diffraction grating. Use unit analysis to obtain the number of mm/line associated with a grating of 600 lines/mm. Enter that value in the data table.

$$\frac{600 \text{ lines}}{\text{mm}} \rightarrow \frac{\text{mm}}{\text{line}}$$

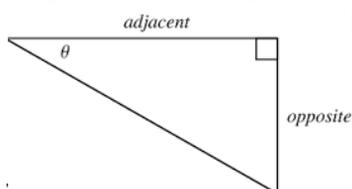
(D.2) Create an Excel spreadsheet to calculate the wavelength of the red, blue-green, and violet hydrogen emission lines from experimental data. Label a block of cells to correspond to the headings below.

(To create the Greek letter θ , go to the *Insert* tab and choose *Symbol*. In the Symbol menu choose the *Greek and Coptic* character set, and *Insert* the theta (θ) symbol. Repeat the process to create the Greek letter lambda, (λ).

Line Color	a (cm)	b (cm)	tan θ	θ (rad)	sin θ	d (nm)	$\lambda_{\text{observed}}$ (nm)
------------	--------	--------	--------------	----------------	--------------	--------	----------------------------------

(D.3) From the data, enter the value for **a**, the distance between the lamp and grating, into the appropriate cell for the red line. (This will be the same for all lines.) Enter the value for **b**, the distance between the red image and the lamp, in the appropriate cell. Enter the value just calculated for **d**, the width of a diffraction grating separation in nanometers.

(D.4) Based on the geometry of the experimental setup, create an Excel formula and calculate the value of tan θ for the red line.



$$\tan\theta = \frac{\text{opposite}}{\text{adjacent}}$$

(D.5) Create formulas to calculate entries for the values of θ and sin θ . Note that the Excel function for \tan^{-1} is called ATAN. Note also that the arguments for all trigonometric functions in Excel are expressed in radians, but you will not need to make any conversions because we do not need the value in degrees.

(D.6) From the entries in the row, create a formula to calculate the experimentally observed wavelength of the red emission line of hydrogen. Note that just employing Eq. 3 will give a wavelength in mm unless you convert the grating separation **d** value to nanometers. As a check, your value for the red line should be within a few percent of 650 nm. If it is not, correct your mistake before proceeding.

(D.7) Copy and Paste your formulas to complete the table for the blue-green and violet lines. Be sure to use the appropriate data for those lines.

(D.8) Create a new block on your spreadsheet to calculate emission wavelengths with the Balmer-Rydberg equation (Eq. 4). Label the cells to correspond to the table below. These terms refer to: the quantum number of the lower state, m ; the quantum number of the upper state, n ; the wavelength, λ ; the frequency, ν ; and the transition energy, E .

(To create a superscript or subscript, type the characters in the cell normally, and then highlight the part to be sub/superscripted. Then *right-click* and choose *Format Cells*, then check the scripting box. Alternately, highlight the pertinent characters and expand the *Font* menu of the *Home* tab. Check the appropriate box.)

m	n	m^2	n^2	$1/\lambda$ (nm^{-1})	λ_{calc} (nm)	ν (s^{-1})	E (J)
-----	-----	-------	-------	----------------------------------	------------------------------	---------------------------	---------

For the Balmer series of emission lines, the quantum number in the lower state is $m = 2$. Because they are emission lines, the quantum number n (in the upper state) must be greater than m . We will calculate wavelengths for specific values of n . By comparison with the observed spectrum, we will be able to associate each observed color with a specific transition; i.e., a specific pair of quantum states.

(D.9) In the first row below the headings, enter the values 2 and 10 in the m and n columns, respectively. Create formulas to calculate the remaining items in the row. The appropriate relations for λ , ν and E can be found in equations (4), (1) and (2), respectively. (Don't succumb to the temptation to do this with your calculator! You will eventually compute over fifty different values in this table. The use of Excel formulas will make this a simple matter, MUCH easier than with a hand calculator. Work smarter, not simply harder!)

(D.10) Use the mouse to select the entire row of numbers just entered or calculated. *Copy* these cells, and *Paste* them into the blank row immediately below. Modify the value of n to calculate the emission wavelength for the $n=9$ to $m=2$ transition of the hydrogen atom. Note that, because of the use of spreadsheet formulas, all calculations were performed immediately when the new value of n was entered.

(D.11) Following the procedure of Step (D.10), complete the table by calculating the emission wavelength for other acceptable Balmer series transitions $m=2$ with $n=8$ through $n=3$. As a check, three of your calculated wavelengths should approximate your three observed wavelengths. If this is not the case, correct your mistake before proceeding.

(D.12) Create a new block on your spreadsheet to compare the three calculated wavelengths to those observed experimentally. Label the block of cells to correspond to the table below. When computing percent error, assume that the calculated value is the accepted or true value.

Line Color	n	m	E (J)	λ_{calc} (nm)	$\lambda_{\text{observed}}$ (nm)	% error
------------	-----	-----	---------	------------------------------	----------------------------------	---------

(D.13) Print out a copy of your spreadsheet, and submit it as part of your lab report. Check for format, labeling, etc. BEFORE printing. For most of this part, 3 or 4 significant figures are appropriate, so adjust your spreadsheet tables to reflect that. All three tables of this part should fit on one page. Make sure that your name and section number appear in the upper right corner of the printout. Note that your laboratory instructor may want an electronic copy of your spreadsheet. Check if this is the case.

Name _____

Section _____

Date _____

PRELAB QUESTIONS
Experiment 40

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 lightbulb?

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

lowest Energy _____ highest Energy

3. Atomic emission from sodium atoms occurs at 589 nm, which causes the distinct yellow color of sodium vapor street lamps. Circle the correct responses below.

a. The emission spectrum of sodium atoms should

i) show a transition at 589 nm

ii) not show a transition at 589 nm

b. The absorption spectrum of sodium atoms should

i) show a transition at 589 nm

ii) not show a transition at 589 nm

c. The transmission spectrum of sodium atoms should

i) show a transition at 589 nm

ii) not show a transition at 589 nm

4. Read the first two pages of the lab manual Appendix I on Spectrophotometry,
<https://www.usna.edu/ChemDept/files/documents/manual/AppendixI.pdf>