

Experiment 40B

WBH 9-3-2022 FV

LIGHT, ENERGY AND SPECTRA

MATERIALS: Hydrogen vapor lamp and power supply (1); Neon vapor lamp and power supply (1); Helium vapor lamp and power supply (1); Quantitative spectrosopes (3); Diffraction grating (1); Rings stands (4); Meter sticks (2); Clamps (4); LED circuit board (6); Leads with Alligator Clips (12); 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 .

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

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

$$E = h\nu = \frac{hc}{\lambda} \quad (2)$$

where the proportionality constant, $h = 6.626 \times 10^{-34} \text{ J s}$, 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. (For a neat hands-on application that demonstrates the principles of additive and subtractive color mixing, see <https://isle.hanover.edu/isle2/Ch06Color/Ch06ColorMixer.html>) Colors absorbed or transmitted are complementary, appearing on opposite sides of a color wheel (see figure 1 on next page). 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.

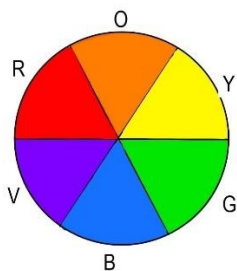


Figure 1. A Color Wheel. Complementary colors appear on opposite sides of the color wheel, and correspond to colors of light absorbed by a sample and those transmitted (or reflected) by it. For example, if red wavelengths of light (R) are selectively absorbed by a solution, then green wavelengths (G) will be transmitted (and scattered) by it. Because our eye detects only the wavelengths of light that are scattered, but not those that are absorbed, we perceive the solution to be green.

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 (figure 2):

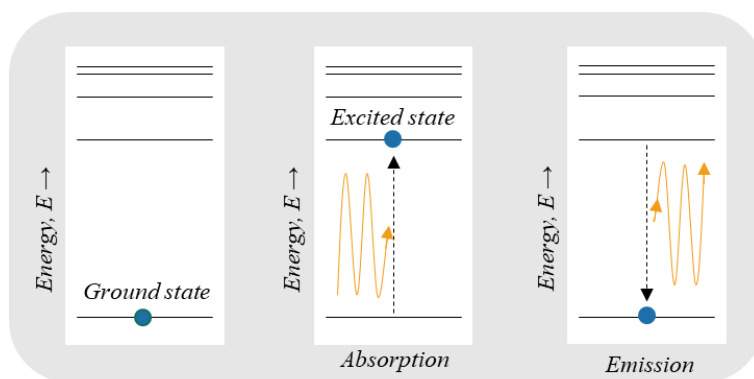


Figure 2. Electron Transitions

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 figure 2, 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. They typically fall into a few categories (table 1):

Table 1.			
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 figure 2). Transmission spectra will show wavelengths *not* absorbed.

Light emission from LEDs also occurs when the system goes from high to low energy, but in this case the energy levels involved are delocalized throughout the solid, rather than being associated with a particular atom in the crystal. These energy levels, known as “bands”, are separated by an energy difference known as the “band gap” (ΔE_g). When sufficient voltage is applied to move electrons from an occupied lower-energy band into an unoccupied, higher energy band, the excited electrons can drop back down into vacancies (or “holes”) in the lower energy band, releasing light as they do so. Because the voltage required to accomplish this and the color (wavelength) of the light emitted both depend on the energy separation between the bands, **the minimum voltage required to light an LED is proportional to its band gap (ΔE_g), and different color LED’s will have different characteristic band gaps.**

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 (or reflective) surface ruled with a large number of parallel, closely-spaced grooves. It is used to separate the light into its constituent wavelengths. If light from an incandescent lamp is directed onto the grating, a continuous spectrum of colors is produced. The grating produces an image of the light source for each color emitted; because all colors are emitted by the glowing filament of the incandescent lamp, these images blur together and appear as a continuous band (like a rainbow). On the other hand, when a hydrogen lamp is viewed through a transparent 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.

As indicated above, the emission spectra of hydrogen atoms and “hydrogen-like” ions (*e.g.* He^+ and Li^{2+} , which also contain only one electron) 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:

$$\frac{1}{\lambda} = R_H \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \quad (3)$$

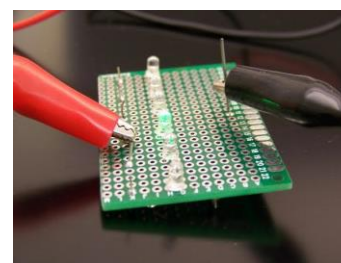
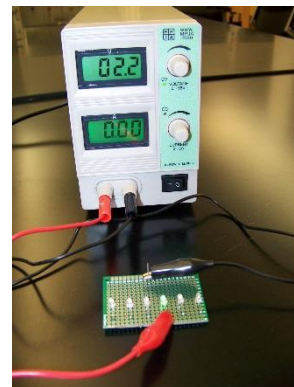
could reproduce the pattern observed by Balmer for hydrogen, where R_H is the Rydberg constant, $R_H = 1.0974 \times 10^7 \text{ m}^{-1}$, and n are positive integers for the states in the emission process. In the Balmer series, $n > 2$. While Rydberg did not understand the physical significance of the n -values, this high regularity and agreement with experiment convinced him that he was on the right track. The wavelengths of the emission lines that you will 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 approx. zero.
2. Observe the polarity (+/-) markings of the two output terminals next to the power switch. When the power supply is connected to a complete circuit, electrons flow out of the negative terminal (which is commonly marked with a black color) and into the positive terminal (which is commonly marked with a red color). Because current can only flow through the LED's in one direction, it is important to connect the wires with the correct polarity, as described below.
3. Using the black wire, connect the negative (black) terminal of the power supply to the single post on the circuit board set off from the row of LEDs. Using the red wire, connect the positive (red) terminal of the power supply to the post just below the right-most LED. See photos at right → .
4. 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. Leave the current knob at this position for the remainder of the experiment.
5. 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 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.
6. 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.
7. Repeat steps 5 and 6 until you have measured the minimum voltages required to excite each of the six LEDs.
8. 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. Using a hand-held spectroscope, observe the emission spectrum of the light bulb. To use the spectroscope, **hold the opening at the narrow end up to your eye while pointing the slit at the left side of the wide end towards the light source.** The "rainbow" spectrum should appear on the screen to the right. A wavelength scale should also be visible on the screen; this scale is calibrated in hundreds of nanometers (i.e.; "6" means 600 nm, etc.) Note: getting your eye, the spectroscope and the source "lined up" properly so you can view the spectrum takes practice, and not all the spectroscopes work equally well, so if you have difficulty, try a different spectroscope!
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. Measurement of the H-atom Emission Spectrum and Comparison with He & Ne

CAUTION: Be sure to **wear goggles** to help screen out ultraviolet radiation given off by the emission tube.

First, you will use a diffraction grating setup to measure the wavelengths of the visible emission lines (Balmer Series) of the hydrogen atom. The apparatus, which allows much more accurate wavelength measurements than can be achieved with the hand-held spectroscopes, will already be assembled in a hood as shown in Figure 3 below:

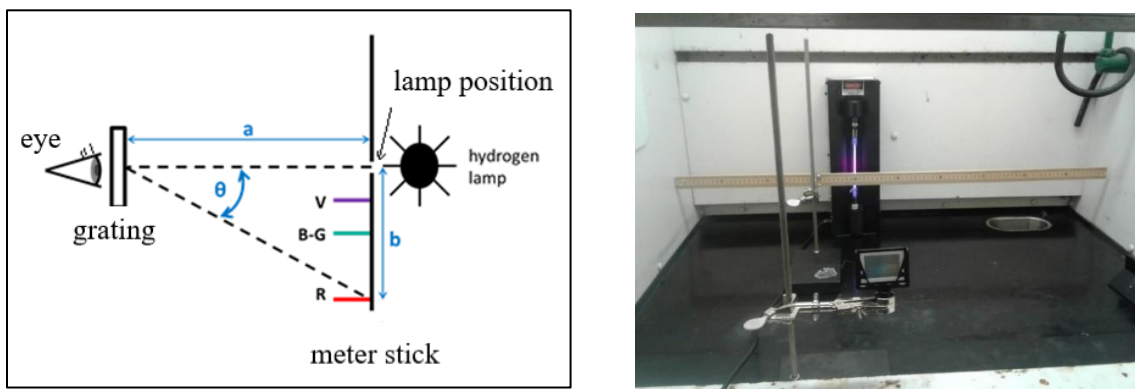


Figure 3. Schematic and picture of the hydrogen discharge lamp apparatus. V, B-G and R stand for violet, blue-green and red and indicate the positions of the diffracted line images of the lamp. From trigonometry, the distances a and b determine the value of the diffraction angle θ , which can be used to calculate the wavelength of each line.

Follow the procedure for measuring the emission line positions as described in the following steps. Record all readings, to the proper number of significant figures and with appropriate units, in the Data Section.

1. Make sure that the grating is parallel to the hood sash and aligned with the lamp, and that the meter stick appears level. The meter stick position of the narrow part of the lamp should be set to 40.0 cm, and the distance from grating to the meter stick (distance a) should be set to 56.0 cm (the correct grating position will be marked with tape).
2. Turn off the hood light and look through the grating towards the glowing hydrogen lamp. You should see three colored images of the emission tube to the right side of the direct line between the diffraction grating and the lamp (figure 4). While looking straight through the grating, note the position on the meter stick where the red line appears to be located. Record this value (to ± 0.1 cm) in the data table provided.
3. Repeat for the Blue-Green and Violet lines.
4. Turn off the hydrogen discharge lamp (unless there is another group waiting to use it).
5. Your instructor will provide an in-lab calculator spreadsheet on the desktop computer in the front of the lab. Input your measured line positions for the three emission lines and record the calculated wavelengths in the data table.



Figure 4. Example of the emission from the hydrogen lamp (obscured by edge of grating) and three diffracted emission lines.

Now, use the hand-held spectroscope to estimate the wavelengths for the brightest emission lines of the helium and neon lamps, which are set up in neighboring hoods:

1. Observe the spectrum of the helium lamp through the hand-held 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.
2. 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 light 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. Measurement of the H-atom Emission Spectrum and Comparison with He & Ne

1. Measured positions of the visible lines in the H-atom emission spectrum, and corresponding wavelength values obtained using the in-lab spreadsheet calculator provided:

Color	Position on meter stick (cm)	Calculated wavelength (nm)
Red (R)		
Blue-Green (B-G)		
Violet (V)		

2. Observations on the emission spectrum of helium discharge lamp as compared to that for hydrogen (similarities, differences, etc.)

Colors and estimated wavelengths of bright features in the helium emission spectrum

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

Colors and estimated wavelengths of bright features in the neon emission spectrum

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 each 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 minimum voltage required to light each diode. The electrical energy 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? Insert values (*with appropriate units*) in the blanks below:

slope: _____ intercept: _____

(A.5) Note by comparison with equation 2 (see page 1), 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 C. Quantitative Determination of the Hydrogen Atom Emission Spectrum

(C.1) In your spreadsheet, enter the wavelength values based on your measurement for the hydrogen lamp emission.

(C.2) 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.3) 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.

QUESTIONS
Experiment 40B

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 electronic states connected by the transition, i.e.

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

Combining $\Delta E = E_{\text{final}} - E_{\text{initial}}$, and $\Delta E = hc/\lambda$ with the above equation, we obtain:

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

where ΔE is the energy of the photon emitted in the transition, h is Planck's constant, c is the speed of light, R_H is the Rydberg constant, and n_{final} and n_{initial} are the quantum numbers of the energy levels involved in the transition. Use the above equation to calculate the value (in Joules) for the following (Balmer series) transition: $n_{\text{initial}} = 6$ to $n_{\text{final}} = 2$. 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 same 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	$1.0974 \times 10^7 \text{ m}^{-1}$		
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 atomic structures 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 552 nm. What is the frequency of this radiation, in Hz?

1b. What is the energy, in joules, of the 552 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 552 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