

Experiment 40B

9.7.2020

LIGHT, ENERGY AND SPECTRA

MATERIALS: Hydrogen vapor lamp and power supply, quantitative spectroscopes (4); LED circuit board; adjustable voltage DC power supply; demo AS-13 flame test kit; incandescent light bulb fixture

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:

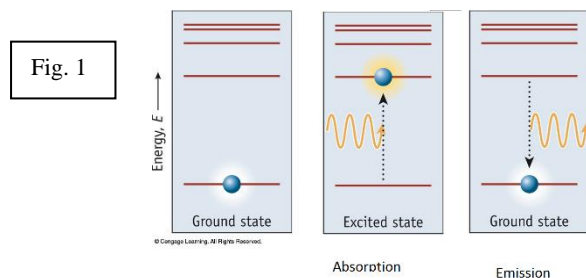
When people hear the term “radiation” they commonly think of *ionizing* radiation associated with nuclear processes. The focus of this lab is a much more common form of radiation: the visible part of the electromagnetic spectrum we call light. Light is a wave phenomenon, characterized by frequency, ν , and wavelength, λ , which are related to each other through the speed of light, c .

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

A scientist named Max Planck showed that the energy of light is related to its frequency by the following:

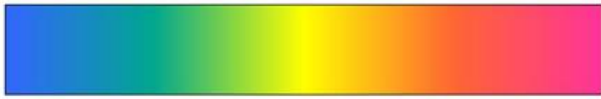

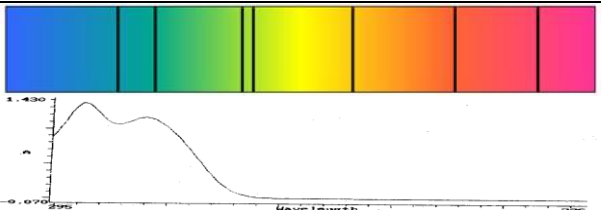
$$E = h\nu = \frac{hc}{\lambda} \quad (2)$$

where the proportionality constant, $h = 6.626 \times 10^{-34} \text{ J s}$. This constant is referred to as “Planck's constant.” Building upon Max Planck's demonstration that the frequency of light is related to its energy, Albert Einstein treated light as a collection of particle-like packets of energy now known as “photons”. A single photon's energy can be quantified using equation 2. 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).



The change between energy states as a result of photons being emitted or absorbed is called a “transition”. 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. Spectra fall into a few basic 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.

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.

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 could reproduce the pattern observed by Balmer for hydrogen:

$$\frac{1}{\lambda} = R \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \quad (3)$$

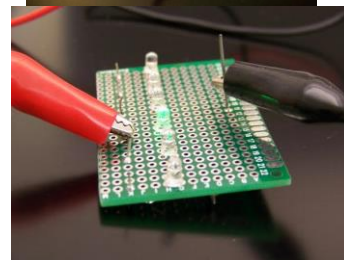
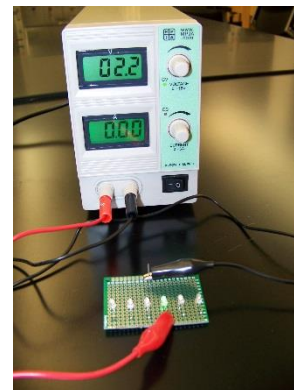
Where n_i is the pre-transition energy level and n_f is the post-transition energy level. R_H is the Rydberg constant, $R_H = 1.0974 \times 10^7 \text{ m}^{-1}$, and n are *positive integers* for the states in the emission process. In the Balmer series $n > 2$. Knowing the relationship, you can calculate 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 in this lab.

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 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 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. Leave the current 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 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 spectrosopes 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.
3. In the instructor hood you will find several small glass dishes containing various ionic solids, and a small amount of methanol. One by one, use the lighter to CAREFULLY ignite the methanol. If there is insufficient methanol to achieve ignition, ask your instructor to refill it. 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. After identifying the color, CAREFULLY put the lid back on the dish to extinguish the flame.

4. 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. Estimate wavelengths based on a spectrum such as Figure 3.1 in your textbook (Gilbert *et al*, 3rd edition).

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: _____

2. Flame test observed emission colors:

Salt	Color	Typical Wavelength (nm)

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, in J per electron, corresponding to each applied voltage. By definition, $1 \text{ V} = 1 \text{ J} / 1\text{C}$, where C is a coulomb, the charge on a single electron, or $1.602 \times 10^{-19} \text{ C}/e$. Therefore, if you multiply the recorded voltage by C, your final answer will be in J/e . 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 $6.63 \times 10^{-34} \text{ J s}$.

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.

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 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 ΔE 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 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 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