Experiment 29C

FV-10/14/11

EMISSION SPECTRUM OF HYDROGEN

MATERIALS: Hydrogen vapor lamp and power supply, diffraction grating (transmission type), meter sticks (2).

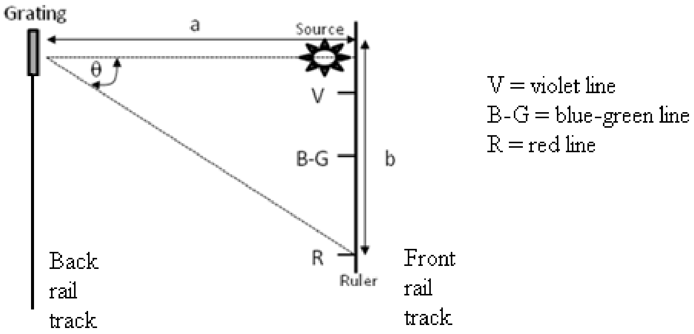
PURPOSE: The purposes of this experiment are: (1) to observe the emission spectrum of atomic hydrogen and (2) to determine the wavelengths and energies of some of the electronic transitions of the Balmer series for hydrogen.

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

1. Explain how a diffraction grating operates.
2. Determine the wavelength of light from its angle of diffraction.
3. Calculate the energy of light from its measured wavelength.
4. Calculate the wavelengths of light expected for specific electronic transitions of hydrogen.
5. Compare the expected wavelengths to the observed wavelengths.
6. Create a simple Excel spreadsheet for the execution of repetitive calculations.

DISCUSSION:

By the middle of the 19th century, scientists knew that atoms were capable of absorbing and emitting visible light. For example, 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. The schematic diagram shows how the wavelengths of the red, blue-green, and violet emission lines of the Balmer series are determined.



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 wavelengths of these emission lines are determined by the diffraction equation

$$\lambda = d \sin \theta \tag{1}$$

where d is the separation between the grooves on the grating, and the angle θ is determined by the geometry of the schematic as shown above. Measurements of the distances shown will provide the angle θ, because $\tan \theta = b/a$.

Note that the geometric relations hold on both the left and right side, as one views the lamp through the grating. Thus the pattern of a continuous band or set of discrete images appears on each side.

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

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

could reproduce the pattern observed by Balmer for hydrogen. R_H is the Rydberg constant, $R_H = 1.09737 \text{ x } 10^7 \text{ m}^{-1}$, and m and n are positive integers. In the Balmer series, m = 2 and n > m. While he 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 (equation 2 if m=2) 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 integers for each transition. Such repetitive calculations are performed easily by computer, using a spreadsheet program such as Excel.

The nature of the light being emitted or absorbed by atoms was a mystery to the physicists who performed the early work in spectroscopy. In 1905, Albert Einstein proposed that light was composed of bundles of energy called photons. Each photon of light contains an energy, E_{photon}, given by the relationship:

$$E_{photon} = h\nu = \frac{hc}{\lambda} \tag{3}$$

where ν is the frequency in s^{-1} and λ is the wavelength in meters. The constants are Planck′s constant, h = 6.6261 x 10^{-34} J·s, and the speed of light, c = 2.9979 x 10^8 m/s. So, once you have determined the wavelengths of your emission lines, you can calculate the energies, E_{photon}, of the photons that are emitted.

But Einstein’s picture of the nature of light occurring in photons did not answer the question as to how the photons of light were produced. In 1913, the Danish physicist, Niels Bohr, proposed the quantum model for the hydrogen atom. In Bohr’s model, the electron can only exist in certain energy levels given by the equation:

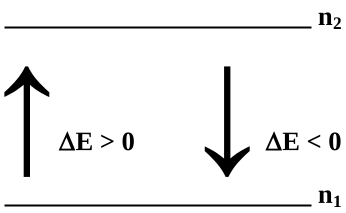
$$E_n = -2.179\ x\,10^{-18}\ J\left(\frac{1}{n^2}\right) \tag{4}$$

where n is an integer known as the quantum number for that particular energy level and 2.17 x 10^{-18} J is the Bohr-Model unit of energy. So when an electron changes energy levels, the difference in energy is given by equation (5).

$$\Delta E = E_\mathrm{f} - E_\mathrm{i} = -2.179\ x\,10^{-18}\ J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \tag{5}$$

Equations (5) and (2) are very similar equations.

The diagram on the right shows the transitions that can occur between two energy states. If the higher energy state is the final state ($n_2 = n_f$) then light is absorbed and ΔE is positive. If the lower state is the final state ($n_1 = n_f$) then light is emitted and ΔE is negative. The magnitude of the ΔE value is equal to the E_{photon} in either case.



Name ______________________________ Section ___________________________

Partner _____________________________ Date ______________________________

PROCEDURE:

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; light from adjacent rooms should be sufficient. Look at the images on the diffraction grating. Three visible lines should be present. Using your line of sight, align the red line emission with the marker on the center of the diffraction grating. Move the needle indicator on the front rail track to be in line with the emission line. Measure the distance of the needle indicator from the lamp.

3. Repeat for the Blue-Green and Violet lines.

DATA SECTION
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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__

Positions of images along front rail track red ____________

blue-green ____________

violet ____________

Distance from front rail track to grating, **a** ____50.0 cm__

DATA TREATMENT

(1) 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, type the letter q in the cell. Highlight the cell and select the q in the text line just above the sheet. Click *Format* from the menu, and choose *Cells*. In the *Font* window, choose *Symbol*. Click *OK*. Repeat the process to create the Greek letter λ, starting with the letter l.)

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

(2) From the data, enter the value for **a**, the distance between the lamp and grating, into the appropriate cell for the red line. Enter the value for **b**, the distance between the red image and the lamp, in the appropriate cell. Enter the value for **d**, the width of a diffraction grating separation in nanometers.

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

you need assistance, see the Excel tutorial at **http://www.chemistry.usna.edu/plebeChem/excel_tutor/index.htm**

(4) From the entries in the row, create a formula to calculate the experimentally observed wavelength of the red emission line of hydrogen. As a check, your value should be within a few percent of 650 nm. If it is not, correct your mistake before proceeding.

(5) Use the mouse to select the entire row of numbers and formulas associated with the red emission line. *Copy* these cells, and *Paste* them into the corresponding cells of the blue-green and violet lines immediately below. Based on the data, modify the values of **b** for the blue-green and violet lines. Your spreadsheet should now include all of your experimentally determined values.

(6) Create a new block on your Excel spreadsheet, to calculate emission wavelengths with the Balmer-Rydberg equation. Label the block of 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, highlight the cell and simply use the mouse to select the characters in the text line just above the sheet. Then click *Format* on the menu. Choose *Cells*, and check the *Superscript* (or *Subscript*) box in the *Effects* field. Click *OK*. To create the Greek letter ν, type the letter u, and proceed as indicated in Step (1) above.)

m	n	m^2	n^2	$1/\lambda$ (nm^{-1})	λ_{calc} (nm)	ν (s^{-1})	E_{photon} (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.

(7) 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. (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 and faster than you could possibly do it with a hand calculator. Work smarter, not harder!)

(8) 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.

(9) Following the procedure of Step (8), 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.

(10) 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	ΔE (J)	λ_{calc} (nm)	$\lambda_{observed}$ (nm)	% error

(11) Print out a copy of your spreadsheet, and submit it as part of your lab report. Make sure that your name and section number appear in the upper right corner of the printout.

(12) Your laboratory instructor may want an electronic copy of your spreadsheet. Check if this is the case.

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

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

2. Equations (2) and (5) are similar equations that calculate different variables corresponding to the emission process. One equation determines the wavelength of the photon that is emitted, while the other equation can be used to calculate the energy of the photon that is emitted. Use equation (3) to show how equation (2) can be converted into equation (5). Calculate the value of the Bohr-Model unit of energy that results from the conversion.

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 constants (we will call them R_{He} and R_{Li}). Knowing the wavelengths, and the appropriate integers for n and m, 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 m and n 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 of m and n as does the red line of H). Record the value of R_{Li} in the table below.

Constant	R (for Hydrogen Atom)	R_{He} (Helium Ion)	R_{Li} (Lithium Ion)
R value	0.01097 nm^{-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 structure</u> of the atoms H, He, and Li? (HINT: focus on the nucleus.)