Matter and Minerals
Fall 2005

Chemistry Lab
Week 8

We will meet in Lab II, 1234 on Thursday of Week 8, from 9 a.m. – 12 noon

"The emission spectrum of atomic hydrogen"

Prepared & Presented by

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Experiment 1: Recording and analyzing the Balmer series of the hydrogen spectrum using the "home made" spectrometer

Do this lab in pairs. You will be given instructions on how to use the "home made" spectrometer.

You have learned about the Balmer series of the emission spectrum of atomic hydrogen in class. Balmer series is in the visible region of the electromagnetic spectrum, hence easy to observe with the naked eye. While it is possible to record the emission spectrum of atomic hydrogen using a state of the art spectrometer, it is easy and more economical to use the "home made" spectrometer.

The hydrogen discharge tube is a quartz tube filled with H_2 gas. Two electrodes (an anode and a cathode) are connected to the ends of the tube. When supplied with an electric charge, there is sufficient energy to break down the bond in the H_2 molecules and to generate excited state H atoms.

$$H_2 \rightarrow 2\ H^* \qquad\qquad H^* = \text{excited state H atoms}$$

These excited state H atoms emit their excess energy, relaxing down to lower energy states. This excess energy is emitted in the form of light. In this experiment, we will be monitoring the <u>visible light</u> emitted by the excited H atoms. We can also monitor UV and infrared light emitted by the excited H atoms provided we have the appropriate detectors.

Visible light is easy to monitor since we can use our eyes as the detector. The function of the diffraction grating is to disperse this emitted light into its respective wavelengths. A grating with approximately 600 grooves/mm is sufficient for this experiment. Gratings with a higher number of grooves/mm provide better resolution, but are more expensive. If we look at the hydrogen discharge tube through the grating, we will see several colored lines; these are the emitted visible light dispersed into different wavelengths by the diffraction grating.

<u>Pre-Lab:</u>

1. Quantum theory predicts that the Rydberg constant

$$R = \frac{me^4}{8\varepsilon_0^2\, c\, h^3} \qquad \textbf{Equation 1}$$

where m = mass of an electron, e = charge of an electron, c = speed of light, h = Planck's constant, ε_0 is the permitivity of a vacuum. Obtain these constant values from standard tables and calculate the value of R using Equation 1. Show all work. Obtain the value of R in cm^{-1} units.

<u>Lab Work:</u>

1. Draw a block diagram of the "home made" spectrometer. Label all components.

2. What kind of a grating did you use in this lab (how many grooves per millimeter)?

3. Record the emission spectrum of mercury in the visible region. Record the positions of the mercury lines in **millimeters** (on the meter ruler) to the highest number of significant figures possible. Record as many Hg lines as possible since this will improve the accuracy of your data.

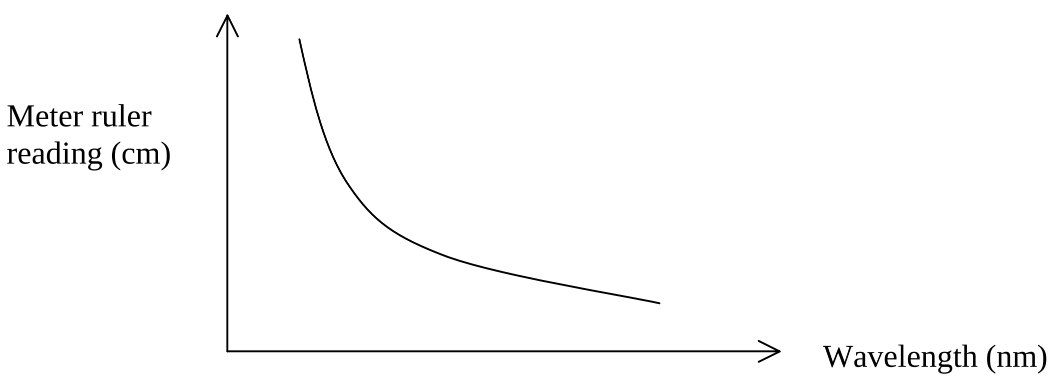
Mercury emission lines in the visible region are at the following wavelengths (nm)

404.6, 435.8, 546.0, 576.9, 579.1, and 695.9

4. Record the emission spectrum of hydrogen in the visible region. Record the positions of the hydrogen lines in **millimeters** (on the meter ruler) to the highest number of significant figures possible. You must be able to see at least **4** lines (red, blue-green, blue-violet and violet). The violet line may not be easy to observe but it will help your data analysis immensely if you spend time to record this line.

Calculations: (Do the calculations alone. You are welcome to get help)

2. Draw a calibration curve (use **Microsoft Excel only**) using the mercury data. Draw a smooth **line of best fit** (this may be a curve rather than a straight line) through the points. See sample calibration curve below.



3. Use your calibration curve to determine the wavelengths of the hydrogen lines.

4. Now you have the wavelengths (λ values) of the hydrogen lines that belong to the Balmer series. You are trying to fit this data to the equation:

$$\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \quad \text{where } n_2 > n_1 \qquad \textbf{Equation 2}$$

where λ is the wavelength, R is the Rydberg constant and n_1, n_2 are principal quantum numbers of the levels involved in the electronic transition. By fitting your data to this equation, you will be able to calculate the value of the Rydberg constant (R) and the correct assignment of the n_2 values. Recall that for the Balmer series n_1=2. The n_2 values for the different Balmer series lines are therefore greater than 2 (since $n_2 > n_1$).

You can rearrange equation Equation 1 to obtain;

$$\frac{1}{\lambda} = -\frac{R}{n_2^2} + \frac{R}{n_1^2} \quad \textbf{Equation 3}$$

Now if we plot a graph of $1/\lambda$ versus $1/(n_2)^2$ we should get a straight line with a negative slope and a positive intercept **provided that we assign the correct sequence of integer values for n_2.**

5. To find the correct sequence of integer values for n_2, make three plots (using Microsoft Excel) for the following three trial sets of n_2 values.
    - n_2 = 3, 4, 5, 6
    - n_2 = 4, 5, 6, 7
    - n_2 = 5, 6, 7, 8

Think carefully about which hydrogen spectral line should be assigned the lowest n_2 value in the above three sequences.

6. **Be sure to have all three plots on the same graph so you can compare them easily.** Do not place the graph on the worksheet. Instead use the "chart" feature in Excel. Your three plots should be labeled appropriately so that the reader can easily understand what you have done. Your worksheet should have your name and labeled columns so that the reader can easily follow your work.

7. When you are satisfied that you have a good graph, print it and the corresponding worksheet and attach to your lab notebook. Label your graphs with your name.

8. Once you determine the correct sequence of integer values for n_2, you can use that sequence for the rest of the data analysis.

9. Using Microsoft Excel, draw a graph of the correct data (i. e. a graph of $1/\lambda$ versus $1/(n_2)^2$ where the n_2 values are now the correct values as determined from the above 3 plots). Using Microsoft Excel, draw a <u>trend line</u> (i.e. line of best fit) through the data points. Display the equation of the trend line on the graph. This equation gives you the slope and the intercept of the trend line.

10. Using the intercept, calculate the value for the Rydberg constant. Could you use the slope of the trend line to calculate the value of the Rydberg constant? Why or why not?

11. Write a short paragraph comparing the calculated value of the Rydberg constant (from the pre-lab) with the value you obtained from this experiment (from 10 above). Suggest reasons for discrepancies (if any) in the two values.

Experiment 2: Emission spectra of various gases

You will be provided with an assortment of gas lamps in the lab. Turn each one on and record the color (naked eye observation) of the gas lamp. Then use a hand-held spectrometer and look at the line spectrum. Record the wavelength and the colors of the spectral lines you observe. Try to be as descriptive as possible (example: **strong** red line at 723 nm, **weak** blue line at 500 nm, **medium** yellow line at 690 nm). **Tabulate your data for easy reading.**

gas lamp	**naked eye observation**	**color of the spectral line**	**intensity**	**wavelength (nm)**