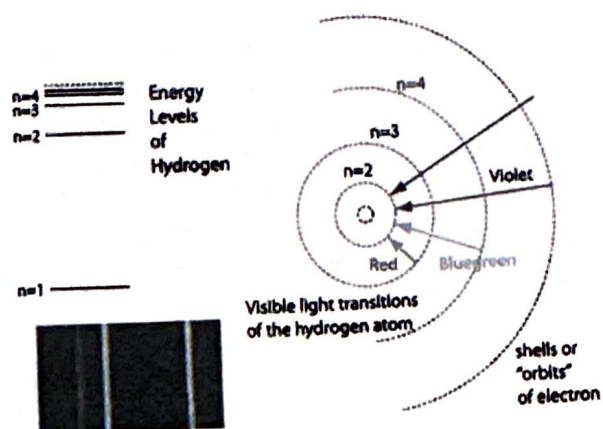# Atomic Spectra – Fingerprinting of Elements

Goals:
- To visually (human optical sensitivity from 400-700nm) verify the atomic spectral lines of Hydrogen in the Balmer Series.
- To successfully apply previously learned diffraction techniques in a new experimental setting.

Motivation:

Analyses of spectral signatures are used in all branches of sciences. These data can help scientist determine the make up of molecules, find minerals on distant planets, and determine reactions in chemistry. Understanding the spectral signatures of elements and molecules is therefore an important endeavor. Although at times these signatures are a complicated mix of elements, in today's lab we will examine the fundamental theory behind such endeavors.

In many laboratories spectral analysis is performed not by one's eyes but by relatively simple to sophisticated laboratory equipment. Spectral lines are seen as peaks in graphs and not colored lines by the eyes. In today's lab we will use a very simple experimental setup with surprisingly precise results.

The first questions to address when viewing spectral lines are: Why are lines seen? Why are these particular lines present for hydrogen? Why is the light not a continuum, a rainbow, as seen with standard room lighting?

The answers rest in the gas itself. Dense gas or solids glow when heated – the rainbow color of the glow is dependent upon the temperature of the dense gas or object. We see this when viewing the dense gas of the Sun, a heating element on a stove, or tungsten wire from an incandescent light bulb.

The tubes that will be used today are filled with hydrogen gas in a dilute concentration and heated to high temperatures adding energy into the system. This is important as the photons (light) emitted during de-excitations can escape without substantial interactions. These photons (light energy) have specific wavelengths (frequencies), which are dependent upon the atomic transitions made by electrons both during excitations (absorption of energy) and de-excitations (releasing energy). These energy transitions are unique to the particular gas and therefore provide, in combination, a fingerprint to identify the gas being observed.

By inspection of the image above, it is noticed that the light color is shown as an arrow next to the shell transition of the electrons in the Bohr model of an atom. The transitions are from n= 3,4,5,6 to the n=2 shell and therefore transitions from a higher excited state to a lower excited state which is not the ground state. Such transitions will take place as electrons, once energetically moved to higher shell levels, will de-excite to a lower shell level almost immediately. Since the electron has no memory, when the electron de-excites (releasing a photon/light) its change in energy may be to any allowable shell level.

1

The Balmer Series equation of hydrogen lines analyzed in this lab was empirically found – this means that the line spectrum was collected and an equation found to fit this data rather than the reverse. The equation shown has been written into a standard format that will allow many hydrogen series to be calculated beyond the visible Balmer series. If one were to change the constant '2' in the equation to '1' the series would describe transitions from higher orbitals to the n=1 state. These electron shell transitions would not be visible, and therefore are not of interest today. It is worth noting that this equation was confirmed by Quantum Mechanics, a branch of physics which examines such atomic transitions.

Balmer's equation:

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{j^2}\right)$$

Here the R is a constant referred to as the Rydberg constant and equal to 1.097×10^7 m^{-1} and 'j' is the energy level the electron occupies before deh exciting to the n=2 shell level.

Method:

In today's experiment you will use a diffraction grating. The equation that governs the diffraction grating (with more than 2 slits) is identical to the double slit experiment you conducted with red laser light previously. To disperse the light into larger angles, so that observations of the different colored lines is possible, the double slit grating will be replaced with a diffraction grating with 600 slits/mm.

Thus, the use of the double slit equation is appropriate and shown below:

$$m\lambda = d\sin(\theta_m)$$

In this equation the m = the order number of the light – if '0' then this is the light color without separation, if m = 1 it is the first time each color is seen on either side on the m = 0 central location, and so on. The 'd' is the slit width – since the diffraction grating has 600slits/mm the distance between each slit must be the inverse of this value. Finally the angle θ relates to the order and the color of the light.

Generally all orders are present simultaneously. For the gratings and wavelength range used today only the orders 0, 1, and 2 will be visible. The data taken is generally for the 1st and part of the 2nd orders. Some spectral lines will be brighter due to the number of energy transitions – the probability of energy transitions, and therefore the

Light Source

Ruler

XL   XR

Smaller λ

θ

L

Larger λ
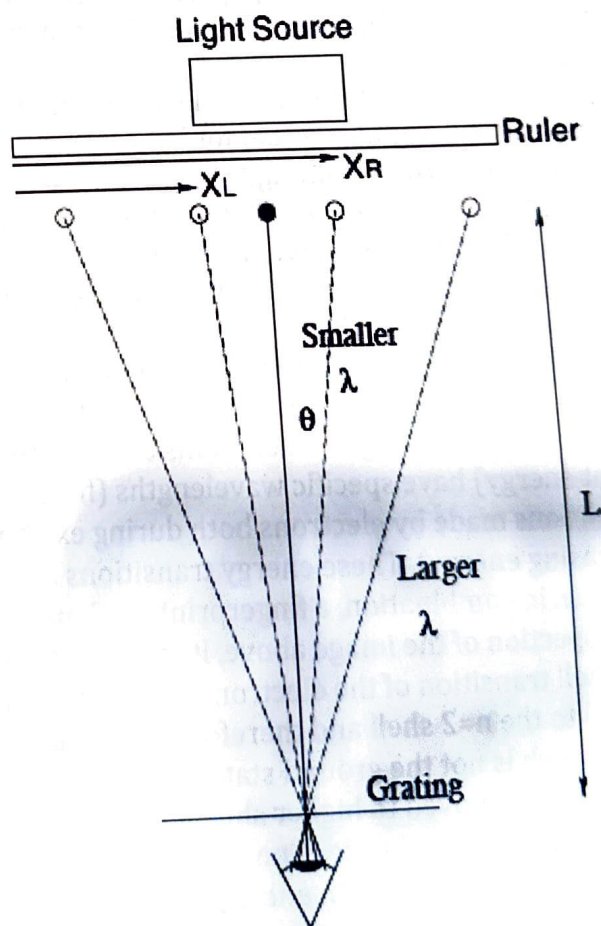
Grating

color of the light seen is governed by Quantum Mechanics.

Viewing the lines will occur by looking directly at a narrow light source through a diffraction grating as shown in the figure on the previous page.

For every order, each discrete color or wavelength, there will be two images – one on either side of the central maxima. A meterstick placed up against the light source will be used to determine the left and right lateral positions. Denoting these values by X_R and X_L, and the source to grating distance of L the angle of the diffraction is determined from the equation

$$\tan \theta = \frac{(x_R - x_L)}{2L}$$

The spectrum lamp is set to face the grating. A meterstick is placed next to it along with a notched black poster board. The optics bench and grating are then placed on boxes directed toward the light such that the hydrogen light that passes through the grating is a bright as possible.

NOTE: The spectrum lamps use high voltage to produce an electrical discharge in the lamp tube. This tube becomes VERY HOT when the light is ON. Exercise caution in handling the lamps and especially do not reach for a tube that has been on.

Procedure:

1. Setup up the equipment as shown above. Look through the grating and ensure everyone can see bands of light. Room lights will need to be turned off. Note that one should not look through the grating a varying angles because this will produce parallax and alter measurements. A flashlight can be used to make measurements on the meterstick. Starting with 'L' (the distance between the grating and the light source) large is a good way to take accurate m = 1 measurements. **The larger L, the larger the angular diffraction of light – why is this statement true? Explain using relevant equations.**

2. With one person holding the flashlight (adjusting so that the one viewing the lines can see both the line and the meterstick – measure the m = 1 first order of the Balmer series – most likely you will see three colors – red, blue-green, and violet blue. There is an additional line – deep violet that not everyone can see. If someone in your group can – make that measurement. (Note the best way to see the deep violet is to focus on the violet blue and see if you can see an additional line in your peripheral vision – the additional line will be at a smaller diffracted angle on either side of m = 0.)

3. When you are done with the first order make sure to measure 'L' and record.

4. To see the 2nd orders (m = 2) you will likely need to move your grating closer to the light source, therefore making the angles smaller for all light bands. Do this (noting that you want 'L' to be small enough but as large as possible) and record as many additional lines as seen. Note - be careful as you may see a m = 3 line before a m = 2 line. **Why could this be? Explain.**

5. Measure the 'L' for the 2nd order and enter all data into Excel. Use the experimentally measured values, and the diffraction equation, to predict the wavelengths of the hydrogen Balmer series.

6. Calculate the Balmer series wavelengths for each color matching the energy shell transitions with the appropriate calculated angle in #5. **Which color has the smallest**

shell transition? What does this mean in terms of change in energy thinking about energy levels of the shells? Explain.

7. Make a plot of $\frac{1}{\lambda}$ vs. $\frac{1}{j^2}$ for both the experimental measured and the Balmer series equation predicted data. Include both m = 1 and m = 2 plots.

8. Add the appropriate trendline to the two plots for the experimentally measured data. **How does the Rydberg constant compare for these plots to the known value? Discuss.**

9. **Thinking about precision and the measurements to determine the angle of diffraction above, state which measured data, m = 1 or m = 2 would have a smaller uncertainty and therefore higher precision.**

10. **Looking at the graph, and your answer for #9, did the data collected demonstrate this expected precision in measurements?**