

Atomic Line Spectrum of Hydrogen

Use of a diffraction grating spectrometer to study discrete spectra

Goals:

- To study the use of a diffraction grating in spectrometry
- To explore the line spectrum of hydrogen

Equipment:

- Spencer spectrometer
- Diffraction grating
- Hydrogen lamp (discharge tube)
- Low power lamp

Introduction:

A **diffraction grating** essentially consists of a large number of very narrow parallel slits arranged next to each other at regular intervals. For a given color of light, a grating will deflect that light from its straight-line path to specific angles which only depend upon the wavelength and the separation of the slits in the grating. At only these very specific angles will any considerable amount of light be transmitted — at essentially every other angle, the transmitted light is negligible. Because the angle to which the light is deflected depends upon the wavelength, a diffraction grating is a very powerful tool to help separate (or disperse) any incident light into its constituent colors (wavelengths).

Specifically, if a parallel beam of monochromatic light is incident on a diffraction grating, we will observe principal maxima (bright spots) at angles θ which obey:

$$d \sin \theta = m \lambda \quad (\text{Eq. 1})$$

where d is the spacing between the slits, m is a whole number known as the **order** of the maximum, and all angles are measured with respect to the central maximum (the straight-line path that all the light would have followed had the grating not been present). If, rather than shining monochromatic light onto the grating, we shine a combination of many wavelengths of light, we can see from Eq. 1 that the grating will deflect each color to a different angle, effectively spreading the incident light into its constituent colors. We thus say that the grating **disperses** the light. A prism could just as easily be used to accomplish this, but good gratings have a much greater dispersive power than prisms. Consequently, they can provide a more detailed picture of the spectra being studied.

In this experiment, you will use a device known as a **diffraction grating spectrometer**, which is simply a diffraction grating mounted onto a very sensitive optical table designed to measure angles quite precisely. With this device, you will endeavor to measure wavelengths of the visible portion of the Balmer series in the line spectrum of hydrogen. Whereas white light is made up of all the colors in the visible spectrum (in other words, it has a continuous spectrum), the emission

or absorption spectrum of atomic gasses are **discrete**: we find that the gas only emits light at very specific wavelengths.

With the photon model of light, we can understand that this discrete spectrum comes about because the electrons in the atom can only occupy discrete energy states. (We would thus say that the energy that an electron can have “orbiting” its atom is *quantized*.) The lowest possible energy state in any atom is what we refer to as the **ground state**. If an atom absorbs a photon with a specific frequency, it can become excited, meaning that the electron would “leap” up to a more energetic quantum energy state. After a short time, if nothing else happens to the atom, the electron will jump back down to a lower energy state, giving up its energy by emitting a photon that carries away exactly the amount of energy lost by the electron. The energy carried by the photon determines its frequency (and also its wavelength) via the relation $E = hf = hc/\lambda$. So, in this picture, atomic gasses have discrete (line) spectra at specific wavelengths and no others because the electrons have discrete energy states.

In the specific case of hydrogen, it was discovered that the allowed energy states are given by the fairly simple relationship

$$E_n = -\frac{13.6}{n^2} \text{ eV}. \quad (\text{Eq. 2})$$

The **Bohr model** of the hydrogen atom helps to provide a theoretical model for understanding why this comes about, but to attempt to calculate the energy levels precisely, we would actually need to invoke the full machinery of quantum mechanics. For our purposes, it is sufficient to utilize Bohr’s model and to note that the energy levels of hydrogen are given by Eq. 2 (and a very similar formula works for hydrogen-like atoms.) The allowed energy states for hydrogen are depicted in an **energy-level diagram** in Figure .

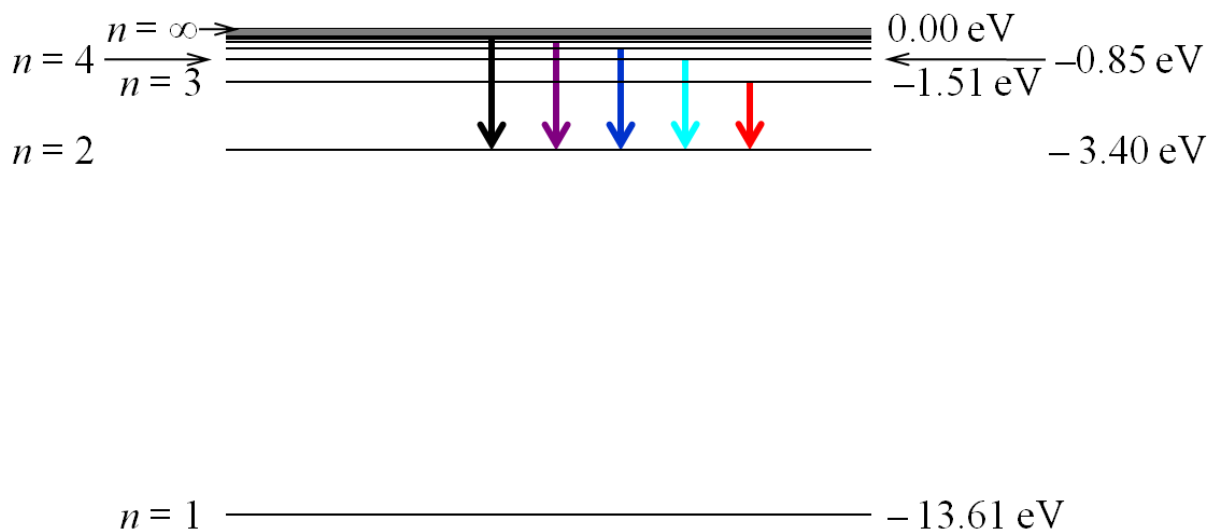


Figure 1 Energy level diagram for atomic hydrogen. The ground state ($n = 1$) is the lowest allowed energy state for the electron. Each excited state has a higher energy (less negative) given by $E_n = (-13.6/n^2) \text{ eV}$. Superimposed on the energy level diagram are depictions of electronic transitions from $n = 3$, $n = 4$, and so on down to the $n = 2$ state. These transitions which occur from higher states down to the $n = 2$ state are known as the **Balmer series**. Only the first four (right-most) of these transitions produce photons which are in the visible range.

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In this experiment, we will excite a gas of hydrogen atoms by applying a high voltage (large potential difference) to the opposite ends of a sealed glass tube containing the hydrogen gas at low pressures (such a device is known as a **discharge tube**). At large enough voltages, the electric field inside the tube becomes sufficient to ionize some of the gas, and the freed electrons can form an electric current through the tube. We can imagine that as these free electrons move through the tube, they frequently collide with other hydrogen atoms, thus imparting some of their energy to the atom and exciting the atom to a higher quantum energy state. As was mentioned above, after a short while, the atom will relax back down to a lower energy state by emitting a photon. So, when we apply a high voltage to this gas discharge tube, we expect it to glow brightly. (This is essentially how neon signs work as well.)

Again, what will be of interest to us here is that the light emanating from the gas discharge tube is not made up of a continuous spectrum like that of the light from an incandescent bulb, but rather a discrete one, made up of only a few wavelengths that are directly related to the quantum energy states of the atom.

Specifically, the four visible lines in the Balmer series of hydrogen have the following characteristics:

Color	Energy of photons [eV]	Frequency [$\times 10^{14}$ Hz]	Wavelength [nm]
red	1.889	4.568	656.3
cyan	2.551	6.167	486.1
blue	2.856	6.906	434.1
violet	3.023	7.309	410.2

In today's experiment, you will attempt to accurately measure these four lines of the hydrogen spectrum.

Name: _____

Sect.: _____

Name: _____

Name: _____

Directions:

The following procedures are best done in the sequence provided. Make certain that any adjustment that is made is optimized before proceeding onto the next step. Correctly following the setup for this activity will yield excellent results.



A: Prism Table

C: Collimator

F: Telescope

H: Eyepiece

L: Divided Circle Clamping Screw

N: Telescope Clamping Screw

B: Prism Table Clamping Screw

D: Collimator Slit

G: Telescope Focusing Ring

K: Divided Circle

M: Divided Circle Fine Adjustment

P: Telescope Fine Adjustment Screw

Activity 1: Preparing the spectrometer

The goal here is to first adjust the eyepiece, then the telescope, and the finally the collimator.

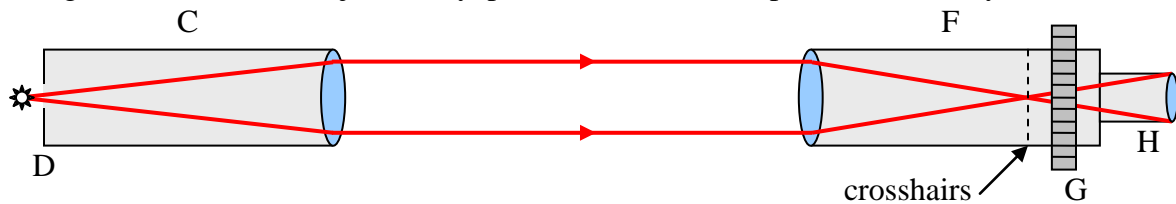


Figure 2 Schematic of the focusing optics

1. Eyepiece (H): Move the eyepiece gently in and out until the crosshairs are clearly visible. (The crosshairs are permanently fixed in the barrel of the telescope.)
2. Telescope (F): By turning the knurled ring (G), focus the telescope on a distant object with a vertical edge. (An object across the room should suffice, but if you are near to a window, a distant building works well.) For final adjustment, move your head left to right and watch for relative motion between the viewed object (the building, for example) and the crosshairs. Apparent motion of this type is called parallax. Keep adjusting the ring until this relative motion disappears. You may need to go back to Step 1 and start over, but the parallax must be eliminated if you are to attain the maximum accuracy.
3. Collimator (C): Illuminate the slit (D) by pointing it outdoors on a bright day, or by using the low-power lamp. Observe the slit through the telescope. Adjust the slit to a width of 1 to 2 mm and bring it into sharp focus by sliding its housing back and forth. Again, move your head slightly back and forth checking for relative motion between the slit and the cross hairs. When this process is completed, orient the slit in the vertical direction.

Reading the Vernier scale

On this instrument there are two windows for reading the angular measurement. You may use either, but once you select a window to use, continue to use it for the entire experiment. The smallest division on the main scale is half of a degree (30 arc-minutes). These are marked on the lower scale. The upper scale, or Vernier, is marked off in 30 divisions, each corresponding to 1 arc-minute. Thus, you'll read the lower scale to the nearest half-degree, and then the Vernier scale will provide you with a 1 arc-minute precision to add on to the reading from the lower scale. Follow along carefully with the example in Figure 3 to learn how to read the scale.

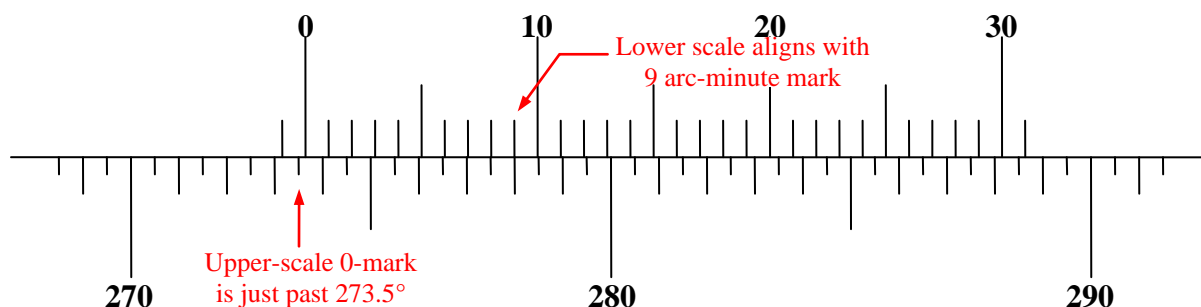
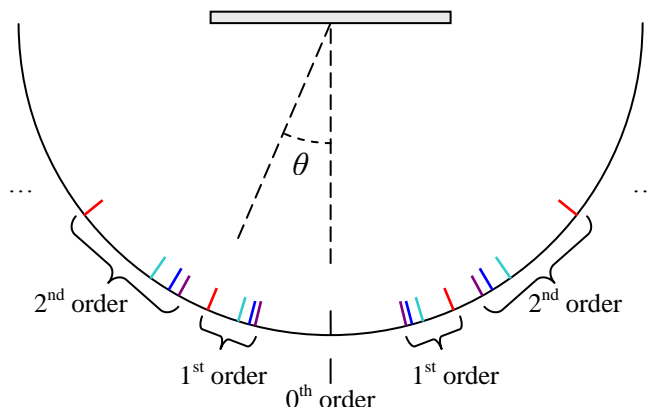


Figure 3 Example of reading the angular position to the nearest arc-minute with a Vernier scale. Use the 0-mark of the Vernier scale to read the lower scale to the nearest half-degree. In this case, the 0-mark is just past the 273.5° mark. Now, in order to read the Vernier scale to the nearest arc-minute, look along the upper scale until you find a mark that lines up directly with *any* mark on the lower scale. In the example above, it is the 9-arc-minute mark that appears to line up precisely with a mark below it. So, to our reading of 273.5° , we should add 9 arc-minutes. Thus, our angular reading is $(273.5 + \frac{9}{60})^\circ$ or equivalently, 273.65° .

4. Record the slit spacing d of the diffraction grating on the next page. Insert the grating with the rulings vertical to the holder on the spectrometer table (A).
5. Unclamp the table (B), rotate the spectrometer table until the grating is perpendicular to the collimator, and re-clamp the table (B).
6. Place the hydrogen discharge tube lamp in front of the collimator slit and turn the lamp on. The lamp has a short life-time, so please turn it off when it is not required.

7. Use the telescope to find the zeroth-order (central) maximum. Record the angular position of this central maximum in the space provided below.



8. From the central maximum, swing your telescope to the right to find the angular position of all the 1st order maxima for the colors in the visible portion of the Balmer series (red, cyan, blue, and violet). For each color, record the angular position in the space provided in the table.
9. Now, find and record the angular positions of all the 1st order maxima on the left-hand side of the central maxima.
10. Find and record the angular positions of all the 2nd order maxima on the left and right of the central maximum.
11. For each color and order, calculate and record the average angular deflection from the central maximum, and from this average angle, the measured wavelength for each color.

Slit spacing of diffraction grating d : _____

Angular position of central maximum (θ_0): _____

Color	Order m	Angular position		$\theta_{\text{avg}} = \frac{1}{2} \theta_L - \theta_0 + \frac{1}{2} \theta_0 - \theta_R $	$\lambda = \frac{d \sin \theta_{\text{avg}}}{m}$	Accepted λ
		Left (θ_L)	Right (θ_R)			

Analysis

Q1. From the table above, determine which of your measured wavelengths most closely matches the accepted value and which differs the most. Report your best and your worst percent differences.

Q2. Using only our eye, we can only see the first four lines of the Balmer series. Which electronic transitions do these emission lines correspond to?

Color	Initial energy level (n_i)	Final energy level (n_f)
red		
cyan		
blue		
violet		

Q3. What is the wavelength of the next line in the Balmer series that is just outside our visible range?

Q4. How would the results of this experiment (that is, the wavelengths you measured) change if you performed the experiment with all of the equipment immersed in water?

Q5. Why could you not find a 3rd order red line for the hydrogen spectrum?
