THE BOHR QUANTUM MODEL

INTRODUCTION

When light from a low-pressure gas is subject to an electric discharge, a discrete line spectrum is emitted. When light from such a low-pressure gas is examined with a spectroscope, it is found to consist of a few bright lines of pure color on a dark background. The wavelength contained in a given line spectrum are characteristic of the particular element emitting the light. Because no two elements emit the same line spectrum, this phenomenon represents a practical and sensitive technique for identifying elements present in unknown samples.

In 1885, by trial and error, Johann Balmer found a formula that predicted the wavelength for the four visible lines in the emission spectrum of hydrogen. The most general form of the formula is given by

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \tag{1}$$

where n_i and n_f are integers, Z is the atomic number (Z=1 for hydrogen), and R is a constant called the Rydberg constant. These four lines of hydrogen are now known as the Balmer series. In 1913, Danish physicist, Niels Bohr proposed a revolutionary quantum model that impacted the scientific community. One of the triumphs of the Bohr Quantum Model was its remarkably successful agreement with the formula proposed by Balmer.

The objective of this lab is:

1. To determine the Rydberg constant by analyzing the Balmer series of hydrogen.
2. To determine the final quantum number n_f for the Balmer series.
3. To "attempt" to determine the initial and final quantum numbers for the electronic transition of Helium for the red, yellow, and purple emission lines by applying the Bohr Quantum Model.

EQUIPMENT

1. Spectrometer (Figure 1)
2. Mercury light source
3. Hydrogen light source
4. Helium light source
5. Black cloth (to cover background light)
6. Diffraction grating
7. Bubble level

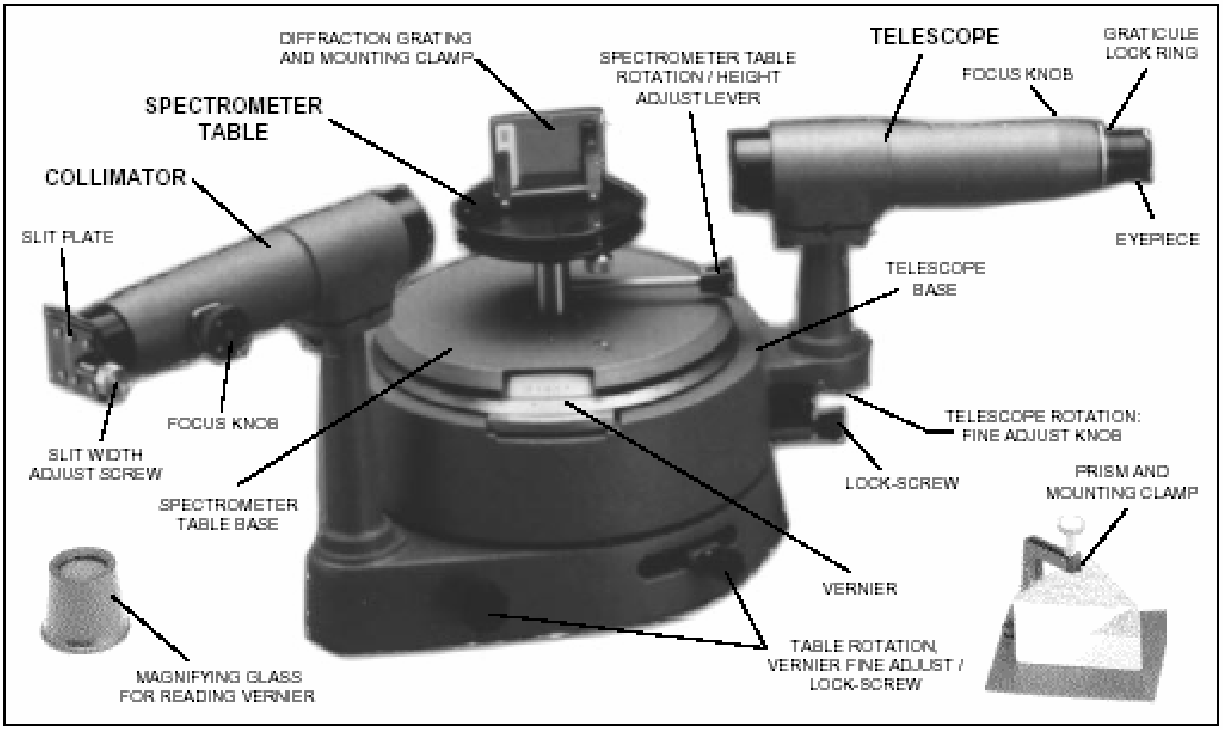

Figure 1 - Spectrometer

THEORY

From your study of diffraction and interference you learned that the condition for interference maxima for a diffraction grating is given by

$$m\lambda = d\sin\theta \tag{2}$$

where m is the order number, λ is the wavelength, and θ is the angle between the mth order and the central maxima.

Bohr derived the expression

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \tag{1}$$

from his quantum model. Initially he derived it for the hydrogen atom but extended it to other atoms in which all but one electron had been removed. Some examples are Li^{2+}, Be^{3+} and He^+.

EQUIPMENT SETUP

Focusing the Spectrometer

1. Level the spectrometer by adjusting the three thumbscrews on the bottom of the table. A bubble level is provided for this purpose.
2. Swing the telescope arm around to roughly 90° from the line of the collimator so that you can view a distant object with the telescope.
3. Point the telescope at a distant object. While looking through the eyepiece, adjust it to give a clear image of the cross hair. Rotate the graticule until one of the crosshairs is vertical.
4. While still looking at the distant object, adjust the position of the objective lens for a clear image of the object in the plane of the crosshairs. There should be no parallax between the object and the crosshairs when proper adjustment is reached. No parallax means that there will be no relative motion between the object and the crosshairs when you move your eye laterally.

Focusing the Collimator

1. Align the telescope directly opposite the collimator.
2. Looking through the telescope, adjust the focus of the collimator and, if necessary, the rotation of the telescope until the slit comes into sharp focus.
3. Tighten the telescope rotation lock-screw, then use the fine adjust knob to align the vertical line of the granicule with the fixed edge of the slit.
4. Adjust the slit width for a clear, bright image.

Measuring Angles of Diffraction

When analyzing a light source, angles of diffraction are measured using the vernier scale. However, the scale only measures the **relative** rotational positions of the telescope and the spectrometer table base. Therefore, before making a measurement, it is important to establish a vernier reading for the undeflected beam. All angles of diffraction are then made relative to this initial reading. See figure 2.

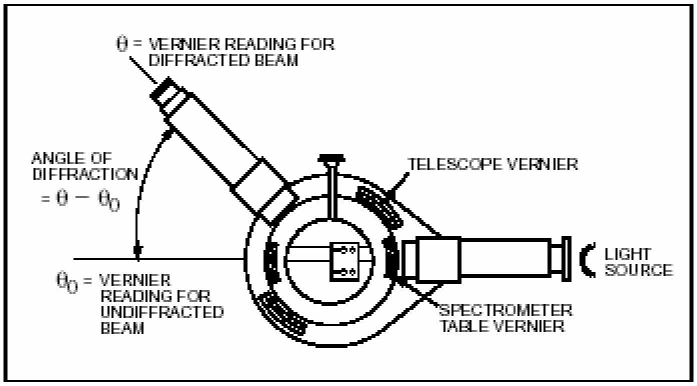

Figure 2

Aligning the Grating

To accurately calculate wavelengths on the basis of diffraction angles, the grating must be perpendicular to the beam of light from the collimator.

1. The telescope must be directly opposite the collimator with the slit in sharp focus and aligned with the vertical cross-hair.

   Perform steps 2-5 with reference to Figure 3.

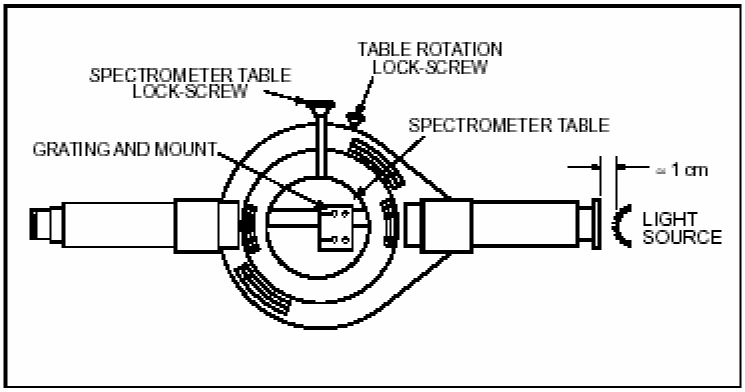

Figure 3

2. Loosen the spectroscope table lock-screw. Align the engraved line on the spectrometer table so that it's collinear with the optical axes of the telescope and the collimator. Tighten the lock-screw.
3. Using the thumbscrews, attach the grating mount so it is perpendicular to the engraved lines.
4. Insert the diffraction grating into the clips of the mount. The grating should spread the spectral lines horizontally.
5. Place the mercury light source approximately 1 cm from the slit. The mercury light sources have a built in green filter to protect your eyes.

   Perform steps 6-9 reference to Figure 4 on the following page.

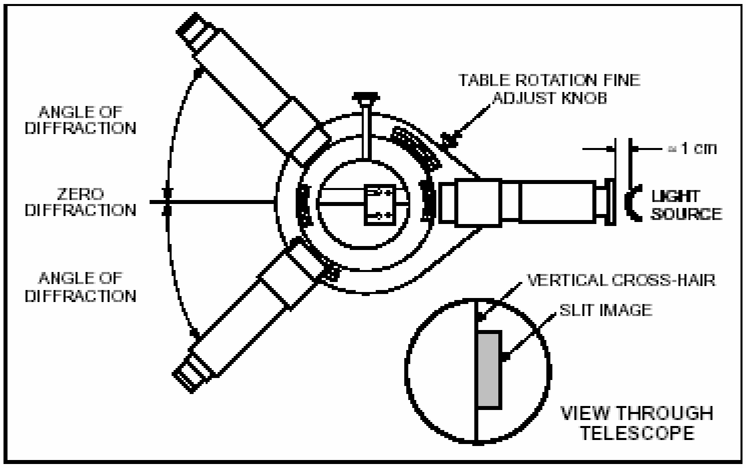

Figure 4

6. Rotate the telescope arm to the right until you see the first spectral line. Align the vertical cross-hair with the fixed edge of the image and carefully measure the angle of diffraction.
7. Rotate the telescope arm to the left until you see the first spectral line. Align the vertical cross-hair with the fixed edge of the image and carefully measure the angle of diffraction.
8. If $\theta_R = \theta_L$ (within 1 minute of arc), then proceed to collect data. If the angles are not equal, then repeat steps 6-8 until they're approximately equal.

(REFER TO LAST PAGE OF LAB TO LEARN HOW TO MAKE THE ANGLE READING WITH THE VERNIER SCALE)

PROCEDURE

Wavelength of Mercury

1. With your spectrometer properly aligned, use the mercury source to obtain data for as many angles θ as you can on both sides of the central maxima (m =0). The wavelength of this green line in the Hg spectrum is 546.1 nm. Compare your calculated average to this value before proceeding to the next section. You should have a small % error.

Spectral Lines of Hydrogen

1. Obtain data for the **first order maxima** on both sides for the spectral lines of hydrogen.
2. The Balmer series spectrum of hydrogen has the same final quantum number n_f for the electronic transition. There should be four emission lines. Compute the wavelength of each spectral line.
3. Correlate each of the spectral lines with a color and initial quantum number n_i. Use your textbook as reference if needed.
4. Using equation (1) from the Bohr Quantum Model plot a graph of $1/\lambda_i$ vs. $1/n_i^2$ such that you get the best linear curve fit.

5. Calculate the Rydberg constant 'R' and also the final quantum number 'n_f' from the equation of the best curve-fit.

Spectral Lines of Helium

1. Obtain data for the **first order maxima** on both sides for the spectral lines of helium.
2. Calculate the wavelengths of the red, yellow, and purple emission lines of helium.
3. For each calculated λ try to determine the initial and final quantum numbers n_i and n_f. This may me accomplished by fitting the data into equation (1a):

$$\frac{1}{\lambda R Z^2} = \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \tag{1a}$$

   For a given λ and Z the left side of this equation is just a constant. Thus, one can obtain, by trial and error, the initial and final quantum numbers n_i and n_f for the electronic transition that gives off the characteristic λ . For example you can set $n_i = 5$ and then vary $n_f = 1,2,3$ and 4 until both sides of the equation are equal. When this occurs you have found the correct quantum numbers! Do this for $n_i \leq 5$ for each of the calculated wavelengths using $Z = 2$.

4. In deriving equation (1a) you used $U(r) = \frac{k(Ze)(-e)}{r}$ for the potential energy function for single-electron atoms. For multiple electron atoms an electron will be subject to the Coulomb attractive force of the nucleus **and** to the Coulomb repulsive force of all other electrons in the atom. These forces will generally cancel each other, leaving a net effective field with potential energy $U_{eff}(r) = \frac{k(Z_{eff}e)(e)}{r}$. In this equation Z_{eff} is the effective atomic number that an electron "sees' and will generally be different for different electrons. The difference in Z_{eff} values for different electrons can be attributed to the shielding or "screening" effects of intervening electrons. For the ground state (n=1) of Helium the ionization energy is 24.581eV and $Z_{eff} = 1.34$. For higher energy states (n >1) $Z_{eff} \approx 1$. Repeat (3) for $Z_{eff} = 1.34$ and $Z_{eff} = 1$.

ANALYSIS FOR HYDROGEN SPECTRUM

1. Compare the Rydberg constant 'R' with the expected value.
2. How does the final quantum number n_f compare with the expected value?
3. Explain any systematic and random errors involved in the experiment and how they effected the outcome.
4. Based on your data did the Bohr Quantum Model successfully describe the Hydrogen spectrum?

ANALYSIS FOR HELIUM SPECTRUM

1. Based on your experimental data for $Z = 2$ can you associate any of the spectral lines with the initial and final quantum numbers n_i and n_f ? Using $Z = 2$ explain why or why not did the Bohr Quantum Model successfully describe the helium spectrum?

2. Based on your experimental data for $Z_{eff} = 1$ can you associate any of the spectral lines with the initial and final quantum numbers n_i and n_f ? Using $Z_{eff} = 1$ explain why or why not did the Bohr Quantum Model successfully describe the helium spectrum?

3. Based on your experimental data for $Z_{eff} = 1.34$ can you associate any of the spectral lines with the initial and final quantum numbers n_i and n_f ? For $Z_{eff} = 1.34$ explain why or why not did the Bohr Quantum Model successfully describe the helium spectrum?

4. If systematic and random errors could be neglected, would you expect the Bohr Model to successfully explain the He spectral lines? How about the spectral lines of other atoms? Give supporting evidence/argument to explain your answer.

REFERENCES:

1. PASCO scientific, Instruction Manual for Model SP-9268 Student Spectrometer, 1991
2. Modern Physics, Serway, Moses, and Moyer, 2nd Ed.

Reading the Vernier Scales

To read the angle, first find where the zero point of the vernier scale aligns with the degree plate and record the value. If the zero point is between two lines, use the smaller value. In Figure 6, below, the zero point on the vernier scale is between the 172° 20′ mark and the 172° 40′ mark on the degree plate, so the recorded value is 172° 20′.

Now use the magnifying glass to find the line on the vernier scale that aligns most closely with any line on the degree plate. In the figure, this is the line corresponding to a measurement of 12' 30" of arc. Add this value to the reading recorded above to get the correct measurement to within 30 seconds of arc: that is, 172° 20' + 12' 30" = 172° 32' 30".