Combined Spectrometry Lab

THE SODIUM D LINES and THE HYDROGEN SPECTRUM

In this lab we will use the D lines in the sodium spectrum to check the calibration of a spectrometer. Then we will use the spectrometer to accurately measure the wavelengths emitted by hydrogen gas. This permits us to test the predictions of the Bohr model of the hydrogen atom.

Note: You should have two diffraction gratings with your setup. They are both delicate and easily damaged. Please handle them only by the edges, without touching the grating part of the surfaces. The more expensive grating (in the orange box) should only be used in the spectrometer. The less expensive grating may be used to look through - but be careful not to damage it.

I. Theory

A. The sodium D lines.

The D lines of the sodium spectrum are very well known. They are a prominent feature of the solar spectrum. They are also used by chemists to identify sodium in a flame spectrum. Yellow flames from a fire are usually radiating the sodium D lines. The accepted values for the wavelength of these two yellow lines are:

$$\lambda_1 = 588.995 \text{ nm}$$
$$\lambda_2 = 589.592 \text{ nm}$$

B. The diffraction grating.

A common way to determine the wavelength ($\lambda$) of light radiated by excited atoms in an emission spectrum is by measuring its deviation by a spectrometer. A grating spectrometer is a device to measure this angle of deviation very accurately. Once the angle of deviation ($\theta$) for an atomic line has been measured, its wavelength can be calculated from the grating equation,

$$m\lambda = d \sin \theta, \quad m = 0, 1, 2, \ldots$$

(1)

where $d$ is the spacing of the slits of the grating. The $m = 1$ line is usually used with a diffraction grating.

C. The Bohr model of the hydrogen atom.

The Bohr model of the hydrogen atom is based on assuming that the electron moves in a circular orbit around the proton, held in its orbit by the electrostatic force. Only those orbits are allowed which have angular momentum equal to an integral multiple of the quantum of angular momentum, $\hbar/2\pi$ (h-bar). It is then possible to predict how much energy is given off as radiation when the electron jumps from one orbit to another. Since the radiation itself is quantized, the
amount of energy given off translates into a particular frequency of light. The wavelength $\lambda$ of
the light given off when the electron jumps from initial orbit $n_i$ to final orbit $n_f$ can be calculated
from the equation

$$\frac{1}{\lambda} = 1.097 \times 10^7 \text{meter}^{-1} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$  \hspace{1cm} (2)

(See Walker, Sec. 31-2.)

The Balmer lines are those wavelengths given off when the final state of the electron has $n_f$=2,
with $n_i$ = 3, 4, 5, 6, 7, . . . . These wavelengths can be calculated from equation (1), giving:

<table>
<thead>
<tr>
<th>$n_i$</th>
<th>$n_f$</th>
<th>$\lambda$ (nm)</th>
<th>color</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>2</td>
<td>656.3</td>
<td>red</td>
</tr>
<tr>
<td>4</td>
<td>2</td>
<td>486.2</td>
<td>aqua-marine</td>
</tr>
<tr>
<td>5</td>
<td>2</td>
<td>434.1</td>
<td>violet</td>
</tr>
<tr>
<td>6</td>
<td>2</td>
<td>410.2</td>
<td>gray (nearly UV!)</td>
</tr>
</tbody>
</table>

Table 1. Balmer-series wavelengths as predicted by the Bohr model.

There are two corrections to these numbers that might be observed in very accurate
measurements. The value of the electron mass used in the calculation should really be the
“reduced mass” of the electron. This is smaller by about 1 part in 2000, and so would increase
the theoretical values of the wavelengths by 1 part in 2000.

The second correction is due to the index of refraction of air; in air, the observed wavelength
should be

$$\lambda_{\text{air}} = \frac{\lambda_{\text{vacuum}}}{n}$$  \hspace{1cm} (3)

For air, $n = 1.00029$, predicting a reduction of the wavelength by about one part in 3000. Since
the reduced-mass effect goes in the opposite direction to the index-of-refraction effect, the net
result is rather small.

II. Experimental Procedure

A. The continuous spectrum.

Ordinary white light from a hot glowing solid or gas is really made up of the spectrum of
colors. If we pass this light through a diffraction grating, interference effects will break up white
light into the familiar rainbow.

1) Pick up the less expensive grating. BE CAREFUL NOT TO TOUCH THE GRATING
WITH YOUR FINGERS!!!

2) Look through the grating at an incandescent source (i.e., the light bulb). What you see is
called a continuous spectrum.

Q1. Sketch and describe what you see.
B. Using the spectrometer to measure the wavelength of light emitted by sodium gas.

Place the diffraction grating so that it is perpendicular to the axis of the slit arm of the spectrometer (the fixed arm). The grating side of the glass should be downstream (facing away from the fixed arm) and over the center of the instrument.

Investigate the motion of the moveable arm. Find out how to read its angle using the vernier scale. Focus the crosshairs by moving the eyepiece in and out. For most accurate measurement, the crosshairs should be aligned so that one is vertical and the other is horizontal.

The spectrometer will work best if the rays of light from the source are parallel when they pass through the diffraction grating. To make this happen, swing the movable arm to a position where you can see a distant object across the room, and focus on it. Now (in the approximation that the object distance is nearly infinite) the incident rays are parallel.

Now, return the moveable arm to the straight-forward position. Turn on the sodium lamp and place it in front of the slit on the fixed arm. Start with the slit closed, then, while looking through the eyepiece, open the slit just enough so that you see a band of light. Adjust the slit so that it is vertical.

Adjust the focusing knob on the fixed arm (not the moveable arm) until the slit is in good focus. Leave the focus of the fixed arm unchanged in this position for the rest of the experiment.

Swing the arm out to the side and find the emission lines of the sodium spectrum. Now how many lines do you see? Roughly sketch the lines, indicating their colors and relative position.

Measure the deviation angle for the yellow doublet of lines (the D lines), measuring to the center of the pair. Record the angles for the left ($\theta_{0L}$) and right ($\theta_{0R}$) deviated images and for the undeviated image in the center ($\theta_0$). Calculate the left and right deviation angles $\theta_l$ and $\theta_r$ for the doublet. They should not be too different. (The “left and right deviation angles” are: $\theta_l =$ angle between positions $\theta_{0L}$ and $\theta_0$; $\theta_r =$ angle between positions $\theta_{0R}$ and $\theta_0$.)

From the equation (1), calculate the wavelengths $\lambda_l$ and $\lambda_r$ for the averaged D lines. Assume that the figure for the number of lines per cm on the grating box is accurate. Then, treating these as two independent measurements, calculate the average value and the error on the average value. Compare this measurement with the average of the accepted values for the two sodium lines, in the standard way (discrepancy in sigmas, quality of agreement).

<table>
<thead>
<tr>
<th>name of line</th>
<th>$\theta_{0L}$</th>
<th>$\theta_{0R}$</th>
<th>$\theta_L$</th>
<th>$\theta_R$</th>
<th>$\lambda_L$</th>
<th>$\lambda_R$</th>
<th>$\lambda_\text{dev}$</th>
<th>$\lambda_\text{accepted}$</th>
<th>discrepancy</th>
<th># of sigs</th>
<th>Q. of Agree.</th>
</tr>
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<tbody>
<tr>
<td>yellow</td>
<td>45.27</td>
<td>88.55</td>
<td>(etc.)</td>
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C. Measuring the hydrogen lines.

Now look at the hydrogen spectrum. You will need to lower the room lights and put a black cloth over the apparatus, because some of the lines are very weak.

How many lines can you find? Going from large angles back towards the central position, you should find a red line, an aqua-colored one (the color of the blue in the Miami Dolphins’ football uniforms), a purple line, and one which is very dim and appears gray to some people. Finding the fourth line is a test of your chances of making a career in the field of spectrometry.

When you have located the lines, measure each of them and determine their wavelengths, following the measurement plan indicated in Table II below.

Compare your results with the values predicted by the Bohr model.

\[ \theta_0 = \underline{________} \quad d = \underline{____________} \]

<table>
<thead>
<tr>
<th>color</th>
<th>(\theta_{0L})</th>
<th>(\theta_{0R})</th>
<th>(\theta_L)</th>
<th>(\theta_R)</th>
<th>(\lambda_L)</th>
<th>(\lambda_R)</th>
<th>(\bar{\lambda})</th>
<th>(\sigma_\lambda)</th>
<th>(\lambda_{\text{accepted}})</th>
<th>discrep</th>
<th># of sigs</th>
<th>Q. of Agree.</th>
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</thead>
<tbody>
<tr>
<td>red</td>
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<td>aqua</td>
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<td>gray</td>
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Table II. EXCEL Data-table format for measurement of the Balmer lines.

III. Equipment

- Precision Tools 5” spectrometer
- Accurate diffraction grating (600 lines/mm)
- Less accurate diffraction grating, for handling
- Pasco sodium lamp (Model SF-9288)
- Discharge tube power supply
- Tensor lamp
- Magnifying glass
- Black cloth