Atomic Emission Spectroscopy

EXPERIMENTAL TASK

To measure the wavelengths of the bright lines in the visible emission spectra of hydrogen and mercury, and calculate the energy of each line.

Objectives

After completing this experiment, the student will be able to:

1. use a spectroscope and discharge lamp to measure the wavelengths of light.
2. calculate energy of light beginning with its wavelength

Additional Reading

- General, Organic and Biological Chemistry, by Timberlake, section 2.6.

Background

The body of knowledge known as chemistry has been built up over thousands of years by people who have had one thing in common: they wanted to be able to explain the universe in terms of cause and effect relationships. Most of these advances in chemistry have developed slowly, and have involved many people taking many small steps. Occasionally, the advances have been major leaps in understanding made by individuals.

One major advance occurred when Niels Bohr, using the first quantum atomic model, explained the cause of the emission spectra of atoms. The concept of the "quantization" of matter in the form of atoms was well accepted early in the 1800's. The idea that energy is quantized, however, was not introduced until 1900. Bohr applied this concept to his

explanation of the hydrogen atomic emission spectrum by stating that electrons in atoms cannot have just any energy, but instead have a few, very specific, "quantized" energy levels available to them.

Bohr's new theory explained the bright-line emission spectra of elements such as hydrogen, where no other theory had been able to do so. When a sample of an element is excited by a high-energy electric discharge, or by heating to high temperature, the atoms emit light. When this light is passed through a prism or diffraction grating, it is spread out into a spectrum, similar to the manner in which water droplets spread sunlight into a rainbow. Unlike the continuous, unbroken spectrum of a rainbow, however, an atomic emission spectrum consists only of a few bright lines of different color.

For instance, the emission spectrum of hydrogen consists of only four visible lines: red, blue-green (or "teal"), violet, and deep violet. The different colors correspond to different wavelengths of light, and therefore to different energies of light. Bohr's atomic theory explains that the different colors, and therefore different energies of light emitted by the atom are due to electrons going from higher principle energy levels to lower principle energy levels. The electrons normally exist in the lowest possible energy level, but can be "excited" to a higher energy level. Then, when the electron drops from the higher energy level to a lower energy level, it gives off energy in the form of light (a photon). The energy of the photon emitted by the atom exactly equals the energy of the higher electron level minus the energy of the lower electron level.

In this experiment you will use a spectroscope and gas discharge lamp to measure the wavelength of each bright line in the visible atomic emission spectra of both hydrogen and mercury. You will then use these measurements (and the equations given later in this handout) to calculate the photon energy for each bright line.

The final results will be the wavelength (in meters) and the photon energy of each bright line measured.

Pre-lab Questions

1. A ladder is said to be 'quantized' because the rungs of the ladder make it so a person may stand at only a few, specific heights. Give 3 examples of other things in daily life that are similarly 'quantized."

2. Use the equations in the "Results" section of this handout to calculate:
   A. the frequency, ν, of light that has wavelength λ = 553 nm.
   B. the energy of a photon that has frequency ν = 3.5 x 10^{15} s^{-1}.

BEFORE STARTING THE EXPERIMENT

Safety

Always keep in mind the rules presented in both the "MCC Laboratory Safety Rules" and the Laboratory Handbook for General Chemistry.   It is your responsibility to make sure that you follow all safety rules at all times, and to graciously help everybody else in the laboratory (including the instructor) to do the same.

The electrical discharge lamp operates at 5000 V! Be very careful to BOTH turn off the switch and unplug the power supply from the outlet before changing lamps. Also, the glass lamps themselves are very expensive (about $100 each), so DO NOT break them.

EXPERIMENT PROCEDURE

Use the spectroscope operation procedure given below to measure the **wavelength** of each line in the bright line spectra of hydrogen and mercury. Construct a data table in your notebook for this information, similar to the table shown below.

Hydrogen		Mercury	
Color of Line	Wavelength	Color of Line	Wavelength

Spectroscope Operation

A number of different spectroscope models are used in the MiraCosta Chemistry labs so your instructor will have to demonstrate the particular model(s) available for your use. The gas discharge lamps will be set up as demonstrated by the instructor, with the overhead lights in the room turned down. Align the spectroscope with the slit toward the lamp making sure that you can see the brightest part of the lamp when you look through the eyepiece--you might have to put books under the lamp holder to make it the right height. Adjust the spectroscope so that you can see the lines of the wavelength measuring scale inside the spectroscope, to the right of the slit. Once everything is adjusted, you will see the spectrum of the element against this illuminated scale. Use this scale to measure the wavelength of each bright line in the spectrum. If the numbers on the scale are "...400...500..." the units for the wavelength are nanometers (nm), but if the numbers are "...4000...5000.." the units are Angstroms (Å). In either case, you will have to convert the wavelength measurements from the units of the spectroscope to **meters** before you can perform the calculations in the Results section.

<u>RESULTS</u>

<u>Using Experimental Values of Wavelength to Calculate Photon Energy</u>
In order to calculate the photon energies, all the wavelength values from the experiment must be converted to units of meters. You should remember the conversion factor for converting nm to m. The conversion for Angstroms is:

$$1 \text{ Å} = 10^{-10} \text{ m}.$$

The energy (E) of a photon is calculated from its frequency (ν), and its frequency is calculated from its wavelength (λ), using these equations:

$$\nu_{(\text{in s}^{-1})} = \frac{2.9979 \times 10^{8} \text{ m/s}}{\lambda_{(m)}} \quad \text{and} \quad E_{(J)} = 6.62608 \times 10^{-34} \text{ J} \cdot \text{s} \times \nu_{(s^{-1})}$$

where "6.62608 x 10^{-34} J s" is Planck's constant and 2.9979 x 10^{8} m/s is the speed of light. Calculate both the frequency and the energy for each bright line in the atomic emission spectra of both hydrogen and mercury.