The Hydrogen Spectrum

In previous laboratory experiment on diffraction, you should have noticed that the light from the mercury discharge tube was composed of only three colors, or three distinct wavelengths of light. The science of spectroscopy was developed around the discovery that each element of the periodic table emits its own characteristic wavelengths of light. The collection of the different distinct wavelengths emitted by an atom is called the **spectrum** of the atom. Spectra which are composed of emitted light are called **emission spectra**, while spectra which are composed of white light with distinct wavelengths which are absorbed or removed are called **absorption spectra**. If one has a collection of several elements, the spectra of the different elements combine or overlap. By comparing the combined spectra to the known spectra of individual elements, one can discover which elements are present. The element helium was first discovered in this manner through the spectroscopic analysis of light from the sun and was only later discovered in material form in natural gas deposits on Earth.

But why *distinct* wavelengths? And why are they different for particular elements. There is nothing distinct about the light from an incandescent source. In an empirical study of the spectrum of hydrogen, Balmer discovered that the precise frequencies and wavelengths of the light produced could be described by an equation involving a constant and an integer. Balmer's equation was then expanded to describe the entire spectrum of hydrogen, including the ultra-violet and the infra-red spectral lines. This equation is called the **Rydberg equation**:

$$\frac{1}{\lambda} = R\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right)$$

where R is the Rydberg constant, and n_1 and n_2 are integers. The presence of integers in this equation created a real problem for physicists until the development of the quantum theory of the atom by Neils Bohr. Bohr's theory suggested that the electron orbiting the nucleus could only have certain *quantized angular momenta*[*]. That means that the electron can only orbit at certain fixed distances and velocities around the nucleus and

[*] It may make it easier to grasp the concept of *quantized* angular momenta if you think of the "Low" "Medium" and "High" settings on an ordinary house fan. Each of these rotational speeds for the fan is an integral multiple (or factor) of the 60 Hz AC electricity which powers it.

subsequently can only possess certain discrete energies. The individual states in which the electron orbits are called **energy levels.** These levels correspond to the integer numbers in the Rydberg equation where n_1 is the **quantum number** of the *initial state* or energy level, and n_2 is the quantum number of the *final state*. Since these energy levels are discrete and quantized, it takes a discrete amount or *quantum* of energy to make the electron move from one level to another, just like it takes a quantized amount of energy for you to walk up one step in a flight of stairs. In the case of the atom, this quantum of energy corresponds to the distinct wavelength (frequency) of light emitted. **Transitions** between different levels product different distinct wavelengths of light. Since the energies of the different levels and the energies of the transitions are determined by the atomic number (the number of protons in the nucleus), each atom has its own characteristic spectrum.

In this experiment, we will be measuring the various wavelengths of the spectral lines of hydrogen, correlating them with their proper quantum numbers, and experimentally determining the Rydberg constant.

Procedure:

Set up the same apparatus as was used last week, but replacing the mercury discharge tube with the hydrogen tube. You will notice four lines in the Balmer series. These are as follows:

Red	656.28 nm
Blue-Green	486.13 nm
Blue	434.05 nm
Violet	410.17 nm

Measure the experimental wavelengths of these four spectral lines using the method from last week, recording their color and wavelength. You will need to measure the wavelength for the first and second-order diffractions of each spectral line.

d = (1/6000) cm = 1666 nm.

Analysis:

The integer numbers in the Rydberg equation are the numbers of each energy level. For emissions in the visible range, the final state (n_2) is level 2. Substituting this into the Rydberg equation gives us the equation for the Balmer series.

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \qquad n = 3, 4, 5 \ldots$$

where the quantum number n is equal to 3, 4, 5... with each larger integer corresponding to a more energetic transition and a shorter wavelength. You will have to associate which value for n goes with each particular spectral line. They should be in order (red=3, blue-green = 4, etc...) but a certain line may be faint and hard to detect.

Substitute the proper measured wavelength and the quantum number to get experimental values for the Rydberg constant. Take caution to get the unit right for R. You should have four values in all, one for each spectral line

.

Average all the experimental values for R together.

Compute the percent error for R. The actual value is R = 109,677.58 1/cm.

Conclusions:

1. Summarize your results for this experiment, reporting your experimental value for R with the percent error.
2. How was the hydrogen spectrum different from the mercury spectrum?
3. Which produces a shorter wavelength, a larger or smaller transition? Why?
4. What do you think the *absorption* spectrum of hydrogen would look like? You may wish to illustrate.

Error Analysis:

What are your primary sources of error? Would a little oxygen or nitrogen contamination in the tube affect the spectrum of hydrogen?

Hydrogen Spectrum

Name:___________________________Section:_________

Abstract:

Data:

Color	*n*	2*x*	*x*	*y*	q	l

Calculations: (Use back if necessary. Show units!)

Calculations:

Calculate the Rydberg constant from each of the wavelengths.

wavelength l	initial state *n*	Rydberg constant R
		R_{ave}

$$\frac{1}{l} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

Conclusions:

Error Analysis: (Compute actual percent errors, and describe sources of error.)