PHYS 1301

The Hydrogen Spectrum

Introduction

In a previous laboratory experiment on diffraction, you noticed that light from a mercury discharge tube was composed of only three colors, or three distinct wavelengths of light. This feature, that an element emits light of specific colors, is an enormously useful probe of how individual atoms of that element behave. Indeed, the science of spectroscopy was developed around the discovery that each element of the periodic table emits light with its own set characteristic wavelengths, or "emission spectrum." of light. If one has a collection of several elements, all emitting light, the spectra of the different elements combine or overlap. By comparing the combined spectra to the known spectra of individual elements, you can discover which elements are present. It is amusing to note that the element helium was first discovered in this manner through the spectroscopic analysis of light from the sun in 1868 and was only *later* discovered in terrestrial minerals in 1895.

But why do we see *distinct* wavelengths in emission spectra? And why are the spectra different for particular elements? There is nothing distinct about the light from an incandescent source such as the ordinary light bulb. In an empirical study of the spectrum of hydrogen, a school teacher Balmer discovered that the precise frequencies and wavelengths of the visible light produced could be described by a simple 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 infrared spectral lines that are not visible to the human eye . 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 Niels Bohr. Bohr's theory suggested that the electron orbiting the nucleus could have only certain *quantized energies*. The implication of this idea is that the electron can orbit only at certain fixed distances and velocities around the nucleus. Individual electron orbits are associated with specific energy levels. Integer numbers uniquely identify these levels and these integers, "quantum numbers," are the ones that show up in the Rydberg equation and that are labeled n_1 and n_2.

The integers in Rydberg's equations identify electron orbits of *specific* radius. In general, the larger the value of the integer, the larger the size of the orbit. Rydberg's equation says that the wavelength of the light emitted from an atom depends on two electron orbits. The interpretation is that an electron makes a transition from the initial orbit identified by the integer n_1 to a final orbit identified by the integer n_2. Furthermore, since there is a *unique* energy associated with each electron orbit, these integers n_1 and n_2 also identify or tag the energy of the electron. Hence, a discrete amount of energy is released or absorbed when an electron makes a transition between two orbits. In the case of the atom, when an electron makes a transition from one orbit to another with a lesser value of its identifying integer, energy is released from the atom and takes the form of emitted light of a distinct wavelength, or equivalently, of distinct frequency.

So the picture we have is that electron transitions between different orbits produce different wavelengths of light and that the actual wavelength value of the light depends on the energy difference between the two orbits. Furthermore, since the energies of the different orbits 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, you will be measure the wavelengths of the spectral lines of hydrogen, correlating them with their proper quantum numbers, and experimentally determine Rydberg's constant.

Equipment

Hydrogen discharge tube**,** 2 x meter sticks, short ruler, diffraction grating & holder.

Procedure

Set up the same apparatus as you did for the diffraction experiment, but replace the mercury discharge tube with the hydrogen tube. You should hopefully be able to see the four lines of the Balmer series:

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

1 nm = 1 x 10^{-9} m. Measure the wavelengths of these four spectral lines using the method from the diffraction laboratory, recording both color and wavelength. Try to make independent measurements of each wavelength from both the first (N = 1) and second-order (N = 2) diffraction of each spectral line. (Note: in this lab, we will use the symbol N to denote diffraction order, so as not to get it confused with the energy-level integers). So the diffraction equation is written

$$N\lambda \;=\; d \sin\,\theta$$

There are 13400 apertures per inch in the grating.

Analysis

The two integer numbers in the Rydberg equation label the orbits, or energy levels, that an electron jumps between when light is emitted. For emissions in the visible range of wavelengths, the final state (n_2) is always level 2. Substituting this into the Rydberg equation gives us the equation for the Balmer series of spectral lines that you observe in this experiment.

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

where the quantum number n is equal to 3, 4, 5..etc.. with each larger integer corresponding to a more energetic transition and a shorter wavelength of light emitted. You will have to associate which value of 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 so you have to be careful to associate each integer with the correct color.

Substitute your measured wavelength and the quantum number to get experimental values for the Rydberg constant. **Take caution to get the correct unit for R**. First ask yourself, if you measure wavelengths in nanometers (nm), what will be the unit of R in the equation above? You should have up to eight R values in all, one for each spectral line and order of diffraction.

Results

Note: use the symbol *N* to denote diffraction order in this lab, so as not to get it confused with the energy-level integer *n*.

Color	***N* (order)**	x	y	θ	λ

Calculate the Rydberg constant from each of the wavelengths and then average .

Wavelength λ	**Initial state *n***	**Rydberg constant R**
		$\mathbf{R_{ave}}$

Discussion

1. Compare your experimental value for R to the actual value R = 109,677.58 cm^{-1} (you will need to carefully convert units).

2. How was the hydrogen spectrum different from the mercury spectrum? Be as specific as possible and say something about the similarity or difference in colors.

3. Which produces a shorter wavelength, a larger or smaller transition?

4. What do you think the *absorption* spectrum of hydrogen would look like? Imagine a rainbow of colors is illuminating a hydrogen tube and you are looking back at the rainbow of colors through the hydrogen lamp. You may want to draw a simple diagram.

5. What are your primary sources of error in determining R? You need to identify these sources and then **explain their significance**.

Conclusions