PHYS 1301

The Hydrogen Spectrum

Introduction

The science of spectroscopy was developed around the discovery that each element emits light with its own set of discrete characteristic wavelengths, or “emission spectrum”. If one has a collection of several elements, all emitting light, 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.

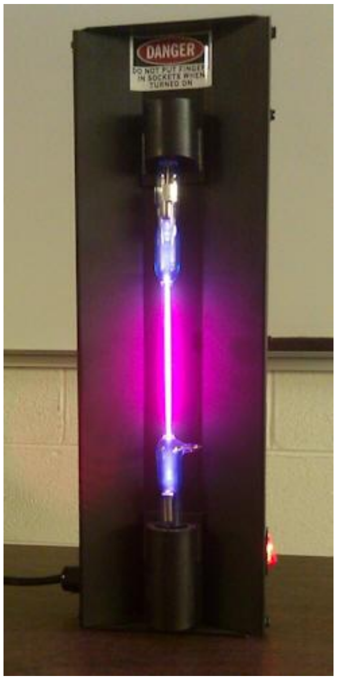
But why do we see *discrete* wavelengths in emission spectra? And why are the spectra different for particular elements? In an empirical study of the spectrum of hydrogen, the simplest atom, a Swiss school teacher Balmer discovered that the precise wavelengths λ of the visible light produced could be 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 light it emits which is 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 discrete integers in this equation created a real problem for physicists until the development of the first quantum theory of the atom by Niels Bohr.

Bohr's theory postulated that the electron inside the atom could have only certain *quantized energies*. Integer numbers uniquely identify these energy levels and these integers, “quantum numbers,” are the ones that show up in the Rydberg equation that are labeled n_1 and n_2. Furthermore, the wavelength λ of the light emitted from an atom depends on two electron energy levels, the interpretation being that an electron makes a transition from the initial level identified by the integer n_1 to a final level identified by the integer n_2. The loss in electron energy E is carried away by one photon of light emitted with the distinct frequency $f = c/\lambda$, via Planck’s formula $E = hf$. So the picture we have is that electron transitions between discrete energy levels in the atom produce different discrete wavelengths of light via the electron energy difference. Furthermore, since the details of the electron energy levels depend on the element (number of electrons, nature of the nucleus, etc.) each atom has its own characteristic spectrum of emitted light.

In this experiment, using the same technique as the laser diffraction experiment, you will calculate the wavelengths of the spectral lines of visible light emitted by hydrogen gas from a discharge lamp using a diffraction grating. The grating will separate the various wavelengths of light that make up of the color of hydrogen because they diffract by different amounts. By correlating these wavelengths with the quantum numbers of electrons transitioning in the atoms, one can experimentally determine Rydberg's constant.



Equipment

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

Warning: the lamp uses high voltage – don't stick your figures near the sockets at the ends of the bulb!

Procedure

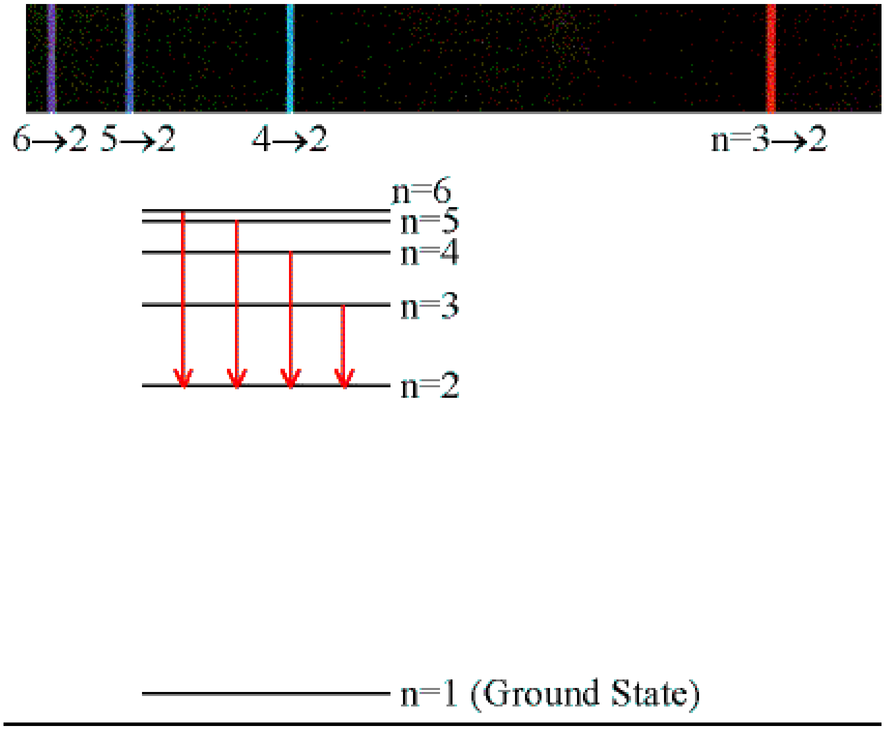
Set up the same apparatus as you did for the laser diffraction experiment, but with the hydrogen lamp at the corner of the L. Looking through the grating straight ahead (N=0), you will see the pinkish glow of the discharge lamp which is a mixture of the following colors and wavelengths

Red	656.28 nm
Cyan (Blue-Green)	486.13 nm
Blue	434.05 nm
Violet	410.17 nm

1 nm = 1 x 10^{-9} m. Looking to one side you will see these colors separated into distinct spectral lines (the so-called Balmer series) corresponding to the N=1 (first–order) diffraction spectrum. The violet line is not very intense and so close to the blue that it may be difficult to distinguish it. Even further to the side, you should also be able to see some and perhaps all of these colors repeated in the N=2 (2^{nd}-order) diffraction pattern. The colors in each order N are the same light emitted by the atom but diffracted by different angles θ through the grating.

Take measurements and calculate the wavelengths of all the spectral lines you can see, using the method from the laser diffraction laboratory, recording color, order and wavelength. Make independent determinations of each wavelength from both the first (N = 1) and second-order (N = 2) diffraction if you see both for a certain color. Here is the diffraction equation again

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



Analysis

The two integer numbers in the Rydberg equation label the energy levels that an electron jumps between when light is emitted. For emissions in the visible range of wavelengths – the Balmer series - the final electron level is always $n_2 = 2$. (Jumps to $n_2 = 1$ yield UV light while jumps to $n_2 > 2$ yield IR light). Substituting this into the Rydberg equation gives us the equation for the Balmer series of spectral lines that you observe in this experiment

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n_1^2}\right) \qquad n_1 = 3, 4, 5, 6$$

You will have to associate which value of n_1 goes with each particular spectral line that you observed. They should be in order, red → 3, cyan → 4, blue → 5, violet → 6, but if a certain line was too hard to detect you have to be careful to skip an integer. Substitute your calculated wavelength and the corresponding quantum number to get a value for the Rydberg constant. You should have up to eight R values in all, one for each spectral line and order of diffraction. Average them and give an uncertainty. **Take caution to get the correct unit for R** - ask yourself, if your wavelengths are in meters for example, what will be the unit of R in the equation above?

Results

Color	**N (order)**	**x**	**Y**	**θ**	**λ**

Calculate the Rydberg constant from each of the spectral lines you saw and then average.

Wavelength λ	**Initial level n_1**	**Rydberg constant R**
		$R_{av} \pm \Delta R$

Questions

1. Which color is diffracted more, red or blue? Hence which wavelengths are diffracted more, short or long?

2. Which produces a shorter wavelength of light, a larger or smaller electron transition?

3. Compare your value for R to the known precise value R = 109,737 cm^{-1}
Is it consistent?

4. Atoms can also *absorb* light by electrons jumping to a higher energy level. Imagine you are observing a rainbow through hydrogen gas. What do you think the *absorption* spectrum of hydrogen would look like? (You may want to sketch it).

5. Explain your primary sources of error in determining R.

Conclusions