

Physics 17 Spring 2003
Lab 3 - The Hydrogen Spectrum

Theory

In the late 1800's, it was known that when a gas is excited by means of an electric discharge and the light emitted is viewed through a diffraction grating, the spectrum observed consists not of a continuous band of light, but of individual lines with well defined wavelengths. Experiments had also shown that the wavelengths of the lines were characteristic of the chemical element emitting the light. They were an atomic fingerprint which somehow resulted from the internal structure of the atom.

In 1885 Johann Balmer discovered a formula which accurately gave the wavelengths of a group of lines emitted by the simplest atom, atomic hydrogen. It was

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

where λ is the wavelength of the emitted light, R is a constant called the Rydberg constant ($R = 1.097 \times 10^7 \text{ m}^{-1}$) and n is an integer greater than 2 ($n = 3, 4, 5, \dots$). This formula, however, was entirely empirical. It could be used to accurately predict the values of the observed wavelengths, but it could not explain why only those specific wavelengths occurred.

In 1913 Niels Bohr succeeded in producing a theory for the structure of the hydrogen atom which explained the sharp line spectrum and gave the proper wavelengths. In so doing, however, he had to reject certain basic classical ideas and replace them with quantum ideas. Like the classical planetary atom, Bohr's hydrogen atom consisted of a nucleus with charge +1 orbited by a single electron with charge -1, but contrary to classical theory the electron was restricted to only certain discrete orbits with energies given by

$$E_n = \frac{E_1}{n^2} = - \frac{13.6 \text{ eV}}{n^2} \quad (2)$$

The lowest energy, the ground state, occurred when $n = 1$. Bohr postulated that while in these allowed orbits (called stationary states), the electrons do not radiate energy even though they are constantly accelerating.

According to the Bohr theory, the hydrogen atom normally is in an unexcited condition with the electron in the ground state. However, when energy is added to the atom (for example, by means of an electrical discharge), the atom can exist in an excited state with the electron being found in a higher energy level. Because the electron has only certain allowed energies, energy can be added to the atom only in discrete amounts. The amount added must be equal to the energy difference between the ground state and the energy of the orbit to which the electron goes. Once the atom has

state by giving off energy in the form of photons. The energy of the photon emitted must again correspond to the energy difference between the two orbital energies. This produces the emission spectrum characteristic of the hydrogen atom.

Suppose, for example, a hydrogen atom originally in the ground state is excited in the electrical discharge to the $n = 3$ state and then relaxes to the $n = 2$ state by emitting a photon. Energy conservation requires that the energy of the photon $h\nu$ be equal to the energy difference between the levels $n = 3$ and $n = 2$,

$$h\nu = E_3 - E_2 = |E_1| \left(\frac{1}{2^2} - \frac{1}{3^2} \right) \quad . \quad (3)$$

If we let $\lambda = c/\nu$, equation (3) can be rewritten in terms of the wavelength of the photon,

$$\frac{1}{\lambda} = \frac{|E_1|}{hc} \left(\frac{1}{2^2} - \frac{1}{3^2} \right) \quad (4)$$

which is consistent with equation (1) if $R = E_1/hc$. When evaluated, this equation gives a wavelength of 6564.7\AA , the observed wavelength of the H_{α} line of the hydrogen spectrum. If, in turn, the 3 in equation (4) is replaced by $n = 4$ and $n = 5$, the expression correctly gives the wavelengths of the H_{β} and H_{γ} lines.

Equation (4) can be generalized by replacing the 2 by n_f (the number of the final energy level) and the 3 by n_i (the number of the initial energy level),

$$\frac{1}{\lambda} = \frac{|E_1|}{hc} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \quad . \quad (5)$$

It can be shown that the constant $|E_1|/hc$ is numerically equal to the Rydberg constant. Making the substitution, we see that equation (5) has the same mathematical form as the Balmer formula and the significance of the Balmer formula becomes clear. It gives the wavelengths of the photons given off when an electron relaxes from a higher energy state to the $n = 2$ state. To obtain the wavelengths of the photons given off when an electron relaxes to the $n = 1$ state, we need only replace n_f by 1 and let $n = 2, 3, 4$, etc.

This set of lines falls in the ultraviolet region and is called the Lyman series. Other series of lines can be found by a similar process. Figure 1 (top of next page) shows some of the transitions for the first three spectral series for hydrogen. Only lines with $n = 2$ in the final state (the Balmer series) are in the visible region of the spectrum and these are the lines we will observe in this experiment.

The Bohr theory is a semiclassical theory that has been found to be of limited value in

understanding the energy levels of more complicated atoms. For the hydrogen atom, however, calculations derived from it give quite accurate results.

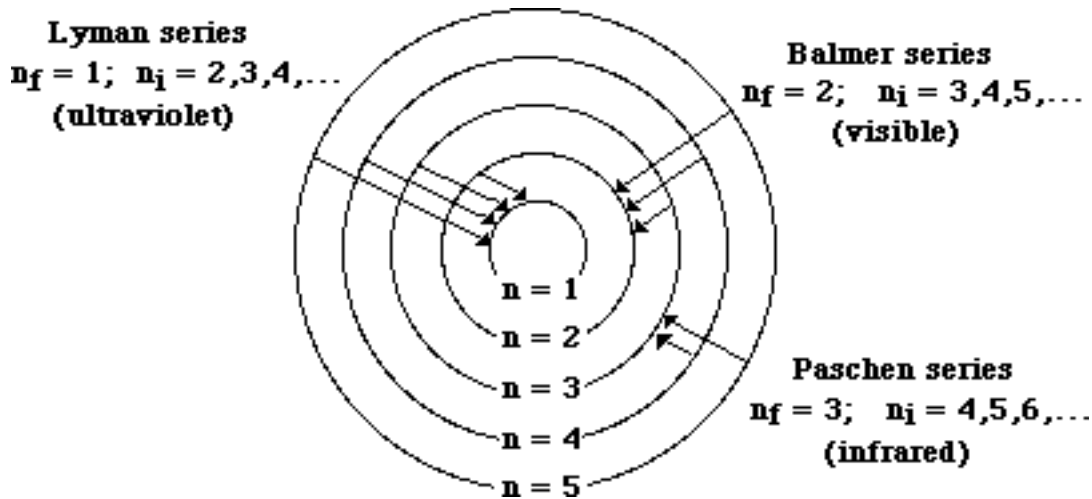


Figure 1

References

Read Beiser chapter 4 pp. 126-136

Experimental Purpose

Bohr's theory of the structure of the hydrogen atom predicts that the energies of the photons of light emitted by excited atomic hydrogen are given by the expression

$$h\nu = |E_1| \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

where h is Planck's constant, ν is the frequency of the emitted light, E_1 is the energy of the ground state of the hydrogen atom, n_i is the number of the initial energy state and n_f is the number of the final energy state.

The purpose of this laboratory is to check the validity of this prediction. This will be done by measuring the wavelengths of the H_α , H_β and H_γ lines of the hydrogen emission spectrum using a grating spectrometer. The measured wavelengths will then be used to compute the energy of the photons producing those lines and those energies will be compared to the energies predicted by the Bohr theory for the hydrogen atom.

Procedure

The apparatus used in this laboratory consists of a hydrogen discharge tube with power supply and a grating spectrometer as shown in figure 2 at the bottom of this page. The hydrogen atoms in the discharge tube are excited by means of an electrical current and the light emitted is passed through the diffraction grating of the spectrometer which disperses it into its component wavelengths. The individual lines can be viewed through the observation arm of the spectrometer. Each line can be precisely located using crosshairs in the eye piece of the observation arm. The angle at which the lines appear is read off a scale on the base of the spectrometer.

The diffraction grating used is a multiple slit system one inch wide with approximately 6000 slits/cm. (The exact value is given on each grating). For any multiple slit system, the angular location θ of the maxima for light of wavelength λ is given by the expression

$$m\lambda = d \sin \theta, \quad m = 0, \pm 1, \pm 2 \dots \quad (6)$$

where d is the distance between the slits and m is the order of the maximum.. In this experiment, we will be concerned only with the first order spectrum ($m = 1$) which simplifies equation (6) to

$$\lambda = d \sin \theta \quad . \quad (7)$$

Once the angle θ for a specific spectral line has been measured, with d known for the grating, the wavelength of the line can be computed from equation (7). When the wavelength is known, the energy of the photons producing that line can be calculated and compared to those predicted by Bohr theory.

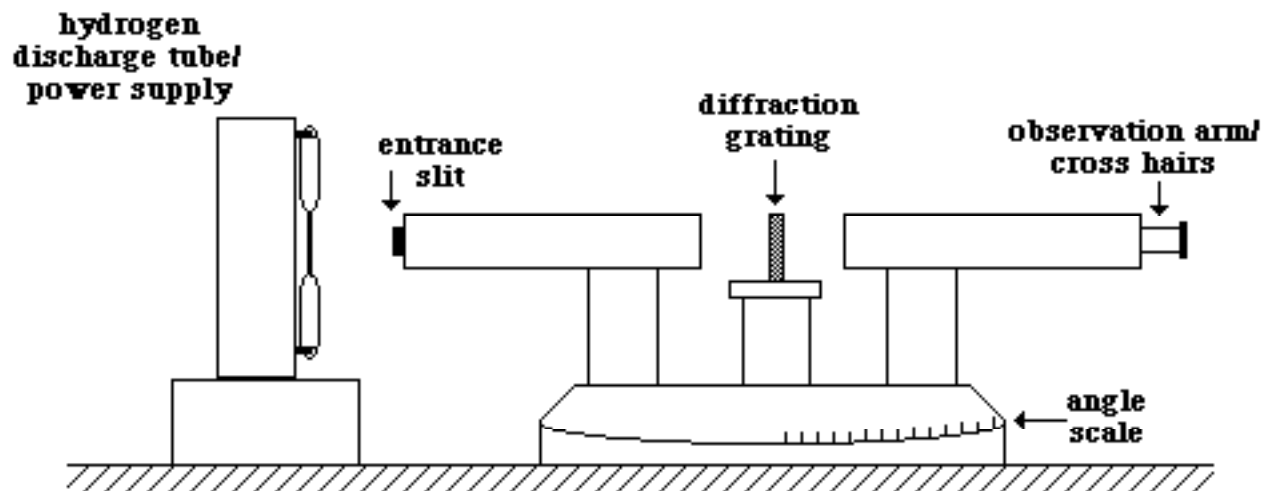


Figure 2

CAUTION: The power supply produces dangerous voltages. Do not touch the metal connectors on the power supply when the power supply is turned on.

1. Make certain the diffraction grating on the spectrometer is level (using three leveling screws under the grating platform) and is in a plane which is perpendicular (both horizontally and

vertically) to the light exiting the arm containing the entrance slits. Do not touch the optical surfaces of the grating. Always handle it by the edges.

Make sure the hydrogen discharge tube is securely in the power supply. Do not touch the thin, light emitting portion of the tube. Plug in the crosshairs illuminator. Turn on the power supply and turn off the lights in the room. Position the thin portion of the discharge tube so that it is just in front of the entrance slit to the spectrometer. Position the observation arm so that it is in line with the arm containing the entrance slit. You should see the bright pink central maximum in the eyepiece.

The thickness of the line can be changed by opening or closing the entrance slit with the knob next to the entrance slit. Adjust the slit opening until a thin line is obtained. The spectrometer optics can be focused by moving the eyepiece in and out in its holder. Focus the spectrometer so that both the observed line and the crosshairs can be clearly seen. [Note: with some of the spectrometers it may be impossible to have both the line and the crosshairs in perfect focus. In that case, focus them so that the crosshairs are clearly visible and the line is as focused as possible.]

2. Swing the observation arm to one side and set the crosshairs on the hydrogen red line (the H_{α} line). There is a fine position adjustment on the lower part of the observation arm which can be used to accurately position the crosshairs. [Note that the line has a finite thickness and the crosshairs can be placed on either edge or in the middle of the line. Where on the line the crosshairs are positioned does not matter as long as you position the crosshairs in the same way each time you make a measurement (why?).] If the crosshairs can not be seen, rotate the crosshair illuminator until just enough light is let into the eyepiece so that both the line and the crosshairs become visible. An alternative method of making the crosshairs visible is to shine a flashlight at an angle on the lens where the light from the grating enters the observation arm. If neither of these methods works or if the line is washed out by the increased light, it may be necessary to open the entrance slit a little more.

The angle scale on the base of the spectrometer is divided into two parts, a fixed part containing the degree markings and a movable vernier scale. Each degree on the fixed scale is divided into two parts each representing thirty minutes of a degree. The vernier scale is divided into thirty equal divisions each representing one minute of a degree. A magnifying glass is provided on the spectrometer to assist you in reading the scales. The scales are read in the following way.

- a. Determine in which degree division on the fixed scale the zero line on the vernier scale is falling. This gives you the whole degree of the angle being measured.
- b. Determine in which half of the degree the zero line on the vernier scale is falling.
- c. Determine which line on the vernier scale exactly lines up with a line on the fixed scale. If the zero on the vernier scale is in the first half of the degree, then the minute measurement is

simply the value of the line on the vernier scale which lined up with the line on the fixed scale. If the zero on the vernier scale is in the second half of the degree, then the minute measurement is the value of the line on the vernier scale which lined up with the line on the fixed scale plus thirty minutes.

Record the angle to the nearest minute.

Repeat this step for the red line on the other side of the central maximum.

Compute λ by finding the difference between the two angles and dividing by 2. This method of measuring λ increases the accuracy the measurement (why?).

3. Repeat step 2 for the hydrogen green line (the H_{β} line) and the hydrogen violet line (the H_{γ} line).
4. Compute λ , ν and E for each line.
5. The H_{β} corresponds to the $n = 3$ to $n = 2$ transition, the H_{γ} to the $n = 4$ to $n = 2$ transition and the H_{δ} to the $n = 5$ to $n = 2$ transition. Compute the theoretical wavelengths and energies as predicted by the Bohr theory and compare them to the experimentally measured values. Compute the percent deviation of the experimental values from the theoretical values.
6. There are discharge tubes containing other gases available in the lab. Observe the spectra of at least two other gases and compare them to the hydrogen spectrum. Choose one gas from each of the following groups:

<u>Group 1</u>	<u>Group 2</u>
Hg	iodine
He	N_2
Ne	H_2O vapor
Ar	

CAUTION: Do not attempt to change the discharge tube without first turning off the high voltage power supply.

After a short period of use, the discharge tube becomes quite hot. Allow a few minutes for it to cool before changing it.

Store the hydrogen tube in its cylindrical storage container while you are looking at the other gases. When you are done looking at a specific gas, please make sure to put it back into its own cylindrical storage container and return to the main storage box.

7. Put the hydrogen spectrum tube back into the power supply leaving the power supply off when you have done so. Unplug the crosshair illuminator.

Lab Report

Follow the usual lab notebook format. Your lab report should include the answers to all of the questions asked in the introduction or procedure, all raw and derived data, and an estimate of the magnitude and sources of error in any data recorded. When answering any question or when giving any comparison or explanation, always refer to specific data to support your statements. For this lab also include the following:

1. a table summarizing all measurements made;
2. all computations made;
3. a table summarizing the final results - theoretical values vs experimental values; and
4. a listing of the similarities and differences noted between the hydrogen spectrum and the other spectra observed, with a discussion of what conclusions you can draw from those observations.