Experiment
21
Spectrum of Atomic Hydrogen

INTRODUCTION

You have no doubt been exposed many times to the Bohr model of the atom. You may have even learned of the connection between this model and bright line spectra emitted by excited gases. In this experiment, you will take a closer look at the relationship between the observed wavelengths in the hydrogen spectrum and the energies involved when electrons undergo transitions between energy levels.

OBJECTIVES

In this experiment, you will

- Use a spectrometer to determine the wavelengths of the emission lines in the visible spectrum of excited hydrogen gas.
- Determine the energies of the photons corresponding to each of these wavelengths.
- Use a modified version of Balmer's equation to relate the photons' energies to specific transitions between energy levels.
- Use your data and the values for the electron transitions to determine a value for Rydberg's constant for hydrogen.

MATERIALS

Vernier LabQuest with LabQuest App **or** computer with Logger *Pro*
Red Tide Emissions Spectrometer
Vernier Spectrum Power Supply and Hydrogen Tube **or** conventional hydrogen gas discharge tube
VIS-NIR Optical Fiber
straight-line filament incandescent bulb and socket
replica diffraction grating

PRE-LAB INVESTIGATION

1. Place the incandescent bulb in the socket, plug it in, and turn on the electrical power to the bulb. View the spectrum through the diffraction grating.

2. Place the hydrogen discharge tube in the power supply socket. Turn on the power and observe the bright line spectrum of hydrogen through the diffraction grating.

Discuss the differences in the spectra from these two sources. The appearance of bright line spectra for excited gases created a seemingly insoluble problem for physicists in the late 19th century. They could not find a simple way to explain why only certain wavelengths were emitted by excited gases. In 1885, Johann Balmer, a Swiss high school mathematics teacher, found an empirical equation

$$\lambda = 364.56\,\text{nm}\left(\frac{m^2}{m^2 - 2^2}\right)$$

where m was an integer greater than 2, that related the wavelengths of the lines in the visible spectrum of hydrogen. Three years later, Johannes Rydberg, a master of spectroscopy, rearranged Balmer's equation and expressed it in a more general form

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

where $1/\lambda$ is the wavenumber (the reciprocal of the wavelength expressed in m), R_H is Rydberg's constant and n_1 and n_2 are integers such that $n_1 < n_2$. This equation led to the discovery of similar sets of spectral lines in the ultraviolet (Lyman series) and infrared (Paschen series) in the early part of the 20th century.

Nevertheless, it was not until Niels Bohr proposed his model of the hydrogen atom in 1911 that a *causal explanation* for the existence of the bright line spectra emerged. Bohr assumed that the electron circled the nucleus in certain well-defined orbits corresponding to specific energy states (see Figure 1 at right). In his model of the hydrogen atom, the electron can exist only in one of these energy states. Ordinarily, the electron exists in its lowest energy condition (called the ground state). So long as the electron is in a particular energy state, the atom does not emit light energy. However, when a hydrogen atom is given enough energy (via an inelastic collision or photon absorption), the electron is bumped up from its ground state (n = 1) to an excited state (n > 1). When the electron drops back to a lower energy state, a photon is usually emitted. The lines in the hydrogen spectrum represent various transitions made by electrons from higher to lower energy states.

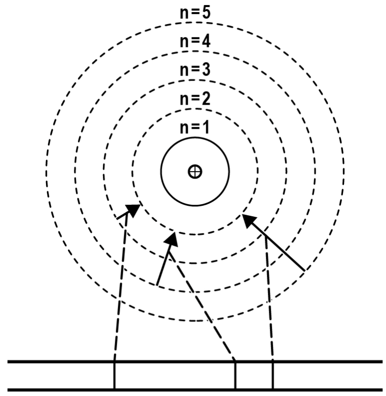

Figure 1

Further detail on the connection between the Bohr model and bright line spectra can by found by observing approximately 5 minutes of a video clip from Episode 3 of Brian Greene's *Fabric of the Cosmos*,[1] which can be viewed from the PBS NOVA web site.

In this experiment you will analyze the visible bright line spectrum for hydrogen and use a variation of the Rydberg equation to relate the energy of the photons associated with each bright line to the energy levels in the Bohr model of the atom.

PROCEDURE

1. If you are using Logger *Pro*, connect the spectrometer to a USB port on the computer and choose New from the File menu. Connect the appropriate optical fiber cable to your spectrometer. If you are using LabQuest as a standalone device, connect the spectrometer to the USB port on LabQuest and choose New from the File menu.

[1] http://www.pbs.org/wgbh/nova/physics/fabric-of-cosmos.html#fabric-quantum. The relevant section starts at time index 6:38 minutes.

2. Set up the data-collection parameters.

   Using Logger *Pro*

   a. Choose Change Units ►Spectrometer ►Intensity from the Experiment menu. The software will measure the intensity in relative units.
   b. Next, choose Set Up Sensors ►Spectrometer from the Experiment menu. Change the data-collection duration to 40 ms.

   Using LabQuest App

   a. On the Meter screen, choose Change Units ► Intensity from the Sensors menu. The software will measure the intensity in relative units.
   b. Change the data-collection duration to 40 ms.

3. Turn on the power to the hydrogen spectrum tube and aim the end of the optical fiber at the middle of the spectrum tube. Your equipment may have a bracket that will hold the end of the optical fiber in position. **Note**: Hydrogen tubes have a limited lifetime, so do not leave the tube on when you are not taking data.

4. Start data collection. If the peak for the red line (H-α) of the spectrum saturates (flat, wide peak at an intensity value of 1.0), move the tip of the optical fiber slightly farther away. If this peak is too small, shift the position of the tip so that more light from the discharge tube enters it. When the intensity of this peak reaches a value of at least 0.8, stop data collection. You need to see at least four distinct peaks to perform this analysis. Depending on your hydrogen tube, you may need to collect two runs: one with no peak height above 0.9, and another with the two strongest peaks saturated at 1.0 so that you can detect the smaller peaks.

5. Once you have at least four peaks visible, save your experiment file. Turn off the hydrogen discharge tube.

EVALUATION OF DATA

1. Use the Examine tool to find the maximum intensity for the red line you observed in the H-spectrum. Record the wavelength. Repeat this step for the remaining peaks.

2. Use the value for the speed of light to calculate the frequency of each of these bright lines. Keep in mind that your wavelengths were measured in nm.

3. Using Planck's equation $E = hf$, calculate the energy of the light emitted for each of the observed lines. Planck's constant is $h = 6.63 \times 10^{-34}\ \mathrm{J \cdot s}$. Record your values in a table like the one below:

Color	Wavelength (nm)	Frequency (1/s)	Δ Energy (J)
red			

4. The photon energies you calculated in Step 3 result from *differences* in the allowed energy states of the electron in hydrogen atoms, $\Delta E = E_{\text{final}} - E_{\text{initial}}$. While they provide a clue to the allowed energy states, these values alone are insufficient to determine the actual energy of the allowed energy states. We can, however, make use of the fact that $\Delta E = hf = hc/\lambda$ to obtain a form of the Rydberg equation that relates the energy of the emitted light to the initial and final energy states.

$$\Delta E = 2.18 \times 10^{-18}\,\mathrm{J}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

5. Your task is to determine values of n_f and n_i that that will give light energies that approximate the values that you calculated in Step 3.One way to approach this task is to divide your calculated values of ΔE by the constant, 2.18×10^{-18} J, then match this fraction by substituting various pairs of integers in the $\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$ portion of the equation. For example if $n_f = 1$, and $n_i = 2$, then the expression

$$\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = \left(\frac{1}{1^2} - \frac{1}{2^2}\right) = 0.75\,.$$

Record your values and best guesses for n_f and n_i in a table like the one below. **Hint:** n_f and n_i for these lines are both less than 8.

Color	ΔE (J)	Ratio ΔE/constant	reasonable values for n_f and n_i
red			

6. Compare your values for n_f and n_i to those obtained by others in your class. Reach a consensus about the values of n_f and n_i for each of the transitions.

7. Determine ΔE for a number of transitions from higher energy levels to the ground state. Determine the wavelengths of the lines associated with each of these transitions. In what region of the electromagnetic spectrum would you expect to find them? Repeat this analysis for $n_f = 3$.

8. Using the wavelengths you obtained in your analysis of the hydrogen spectrum and the values of n_f and n_i corresponding to the transitions producing these lines, determine an average value for the Rydberg constant for hydrogen. The Rydberg equation is in the Pre-Lab Investigation. How does your experimental value compare to the accepted value for this constant?

EXTENSION

Visit the Visual Quantum Mechanics web site, http://phys.educ.ksu.edu/vqm/free/h2spec.html, and run the simulation for hydrogen spectroscopy. Note that a hydrogen gas discharge tube is lit and a spectrum appears at the top. Step-by-step directions are provided. See how closely your simulation spectrum matches the observed spectrum for hydrogen.

Vernier Lab Safety Instructions Disclaimer

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This copy does not include:

- **Essential instructor information including discussions on how to lead students to a successful activity**
- **Sample data**
- **Important tips for successfully doing these labs**
- **Answers to questions and extensions**

The complete *Advanced Physics with Vernier – Beyond Mechanics* lab manual includes 22 experiments, as well as essential teacher information. The full lab book is available for purchase at: www.vernier.com/phys-abm

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