Introduction

In this lab, you will use a diffraction grating to split up light into the various colors which make up the different wavelengths of the visible electromagnetic spectrum. You will assemble a spectrometer, incorporating the diffraction grating. A spectrometer is a device that allows you to separate and view the colors present in a light source and to measure the wavelength for each color. You will then examine light from a variety of different discharge tubes containing gases at low pressure. These emit line spectra, where light is present only at certain wavelengths (or colors). The wavelengths in a line spectrum arise from transitions between energy levels of the allowed states of electrons in the atoms. By comparison of the spectrum you observe with that on the chart provided, you will be able to identify the gas in the discharge tube. At the end of the lab, the spectrometer is yours to take home.

Before you start, review diffraction and the function of a diffraction grating (Serway, Section 24.6) and the Bohr model of the Hydrogen atom in Chapter 28. **IT IS IMPERATIVE THAT YOU READ CHAPTER 28 AND SECTION 24.6 BEFORE COMING TO LAB!**


Pre-Lab Homework:

1. The grating equation tells you at what angles different colors of light will appear:

$$d \sin \theta = m\lambda$$

where d is the separation between the lines on the grating,
  m is an integer (i.e., $m = 0, \pm1, \pm2, \pm3$, etc.) known as the diffraction order number,
  λ is the wavelength (with the same length units as d), and
  θ is the angle at which light with wavelength λ is diffracted.

Make a sketch based on the figure below to show a narrow parallel beam of white light shining onto a grating with line separation $d = 1000$ nm. Show the angles for red and blue light for $m = 0, +1$, and -1 by drawing them in appropriate colors. You can find the approximate wavelengths for red and blue light in Chapter 28.


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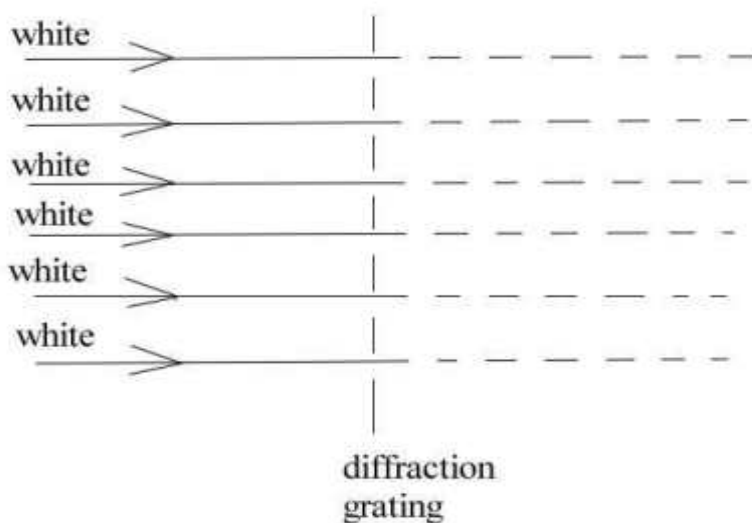
2. From the Bohr model of the Hydrogen atom, calculate the minimum amount of energy (in eV) an electron in the lowest orbital ($n = 1$) would need to free it from its proton (*i.e.*, to ionize the atom). Also, calculate the calculate the minimum amount of energy (in eV) an electron in the second lowest orbital ($n = 2$) would need to free it from its proton.

3. In this lab you will examine the spectral lines of the Hydrogen atom known as the Balmer series. Another series for Hydrogen is called the Lyman series and has an equation of the form:

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

Calculate the first four wavelengths (use the n values listed above two through five) of the Lyman series. Where R is the Rydberg constant. Are these four wavelengths in the visible range of the spectrum?

Experiment A: Verify the Spectrometer

The spectrometer restricts your view to the $m = 0$ and $m = +1$ orders of the diffraction pattern. Practice holding your spectrometer at the correct viewing angle. There is a slit on the wide end of the spectrometer (as shown in the picture to the right). *Aim the spectrometer's slit at the light source to allow light to enter the spectrometer*. You should be able to view the colors along the scale printed on the film by looking through the eyepiece.

Eyepiece

Light goes in here
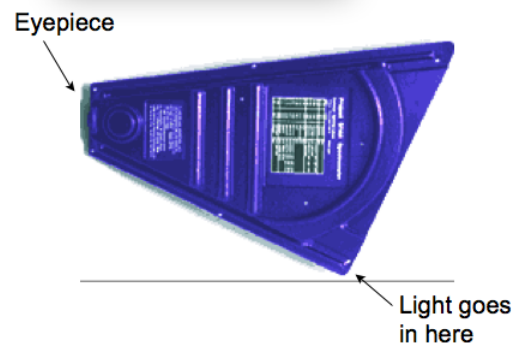
Look at the scale on the spectrometer. The top units are electron volts (eV). This is simply an alternate unit of energy. The conversion factor to Joules is 1 eV = 1.6×10^{-19} J. This energy unit is used frequently for the energy electrons have in atoms. The bottom scale is in nanometers (1 nm = 10^{-9} m) and corresponds to the wavelength of the light. Confirm that you see red light at approximately 650 nm and blue light at approximately 450 nm.

Before you use the spectrometer to precisely measure wavelengths (and hence identify elements), you must confirm that the wavelength scale is correctly installed. To do this, use the bright green line that is present in the spectrum of the fluorescent ceiling lights. This line is known to have a wavelength of 546 nm, as written on the top of the spectrometer. (In fact, there is an additional hash mark on the wavelength scale corresponding to exactly this wavelength.) If necessary, you can slide the film scale to the left or right in its curved slot to adjust the calibration. If the emission lines don't appear to fall on the film scale, you can also rotate the grating lens in its holder to adjust what you see. Check that you see this line in the correct place on the scale. If you are unsure if you have properly set up your spectrometer, check with your TA or LTAC.

Experiment B: Measuring the Rydberg Constant

In this part of the lab you will use the glass discharge tubes. These contain gas at low pressure. The gases are excited by a high voltage supply. Please read carefully the warning below before using the discharge tubes. If you are at all uncertain or uncomfortable about how to run the tubes, get help from your TA.

> **WARNING: Discharge tubes are powered by very high voltages (up to several thousand Volts) that can be lethal. Keep your fingers and jewelry away from the ends of the tube and the sunken electrodes. Before changing tubes, always switch off the power supply and physically unplug it from the socket. Hold the tube in the middle, away from either end, with a paper towel (since the tube can be extremely hot). Always use only one hand when working with these tubes and the power supply to minimize the danger of an electrical shock. If you use two hands, the current could pass from one hand to the other, via the heart, which could be fatal.** *Always turn the power supply off when not in use.*

First, obtain a hydrogen discharge tube and place it (carefully!) in the holder. If you have difficulty holding the spectrometer so it's slit points at the light, you may rest the spectrometer on the wooden stand at the end of your lab table. Adjust the height of the spectrometer to center the slit with respect to the hydrogen tube. Set up your table lamp behind the tube facing the wall, to provide enough light to see the emission scale on your film. Look through the grating and locate the spectrum of the hydrogen diffracted into the $m = 1$ (first) order. Compare the spectrum you see with the color chart on the wall. Can you see the H_δ line in the deep violet, the faintest of the four lines in the visible spectrum?

Measure and record on the chart below the wavelengths of all the lines you can see. *Remember, your ability to distinguish between colors isn't important, many people see colors in slightly different ways. The placement of the emission line on the proper wavelength location on the graph is what identifies the element.* Compare your chart with the table. Is your spectrometer a reliable calibrated instrument?
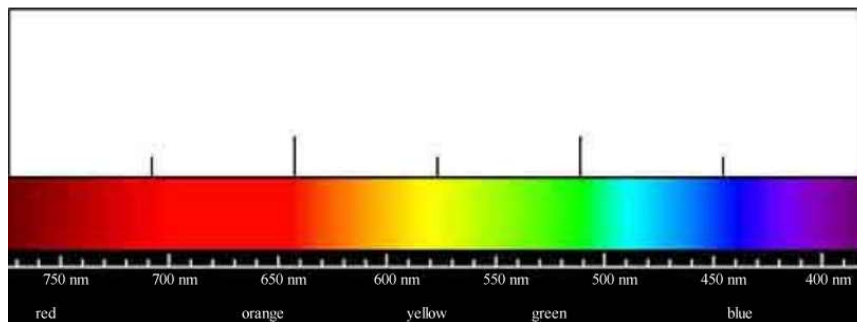

Choose one of the Hydrogen lines from your measured values. Identify its quantum number, n, from the table. Use Balmer's equation,

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

to calculate the value of the Rydberg constant, *R*, from your measured value of the wavelength of your chosen spectral line. Repeat this calculation with a different line. How do the values you calculated compare to the accepted value of $R = 1.1 \times 10^7$ m^{-1}. What is the significance of *R*?

Now pick two additional tubes of known elements and record their emission spectra in a similar fashion. Record their spectra on the following charts. For each element label at least three lines that you can identify from the table.
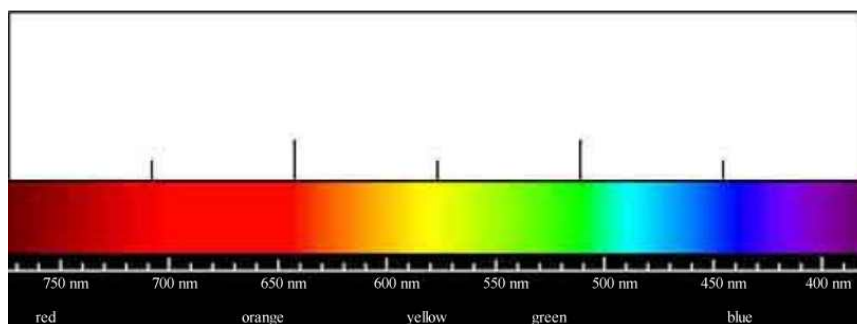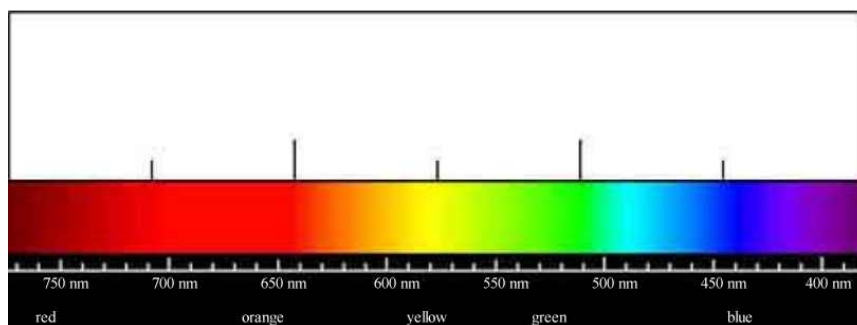

Experiment C: Identifying an Unknown Element

View at least two of the unknown gas discharge tubes with your spectrometer. To identify the elements present in each light source, you may use the printed table at the end of the lab manual, or go to the following website: http://jersey.uoregon.edu/vlab/elements/Elements.html. Visually compare the spectrum you see to the spectrum in the given website. If the computer monitor casts too much background light while you are viewing the tube, put the computer monitor to sleep temporarily while you make your measurements. (Ask your TA how to do this if you are unsure; please do not turn the computer off). Record both the label and your identification of each tube.

Experiment D: Identifying Other Light Sources

Your TA will perform several experiments for you at the front of the lab. However, you must use your own spectrometer to observe the results. Be careful when near the propane torch!

D1. The TA will burn some sodium for you, creating a visibly yellow flame. View the burning salt with your spectrometer. Can you see any sodium emission lines? Measure and record their color and wavelengths on the chart below.
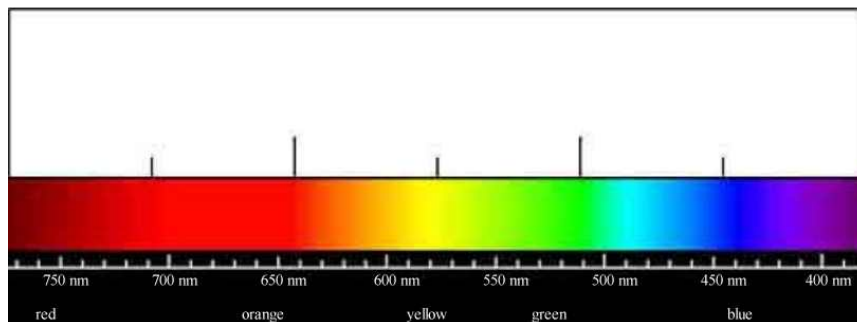

D3. All objects emit *thermal radiation.* The spectrum and amount of energy emitted depend on the object's temperature and the surface properties. The hot filament in a light bulb emits thermal radiation as both infrared radiation and visible light. The sun emits its energy primarily in the visible part of the electromagnetic spectrum. The Wien Displacement Law describes how the wavelength at which the most radiation is emitted varies as the temperature changes.

$$\lambda\,T = 0.03 \text{ m.K}$$

Here λ is the wavelength at the peak of the spectrum, and T is the temperature in degrees Kelvin. The figure shows the thermal radiation for objects at 2000K, 3000K and 4000K. The dotted lines show the wavelengths of blue and red light. We can see only the radiation between these wavelengths.


Serway/Jewett; Principles of Physics, 3/e Figure 28.2
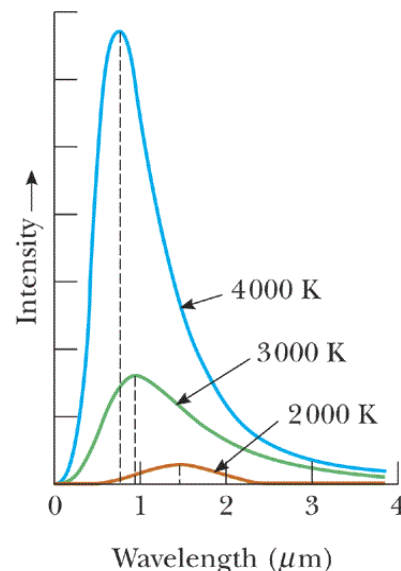
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Look at the spectrum of the incandescent light bulb attached to the variable voltage supply. The TA will start at a low voltage and slowly increase it. What are the effects of increasing the voltage on the spectral distribution you see with your spectrometer? At low voltage, what wavelengths (or what color) appear the brightest? Does this change as you increase the voltage? Why does this happen?

D4. Once again, view the fluorescent lights on the ceiling of the lab through the spectrometer. Graph the visible emission lines on the blank graph below. Can you identify what is inside these lights?
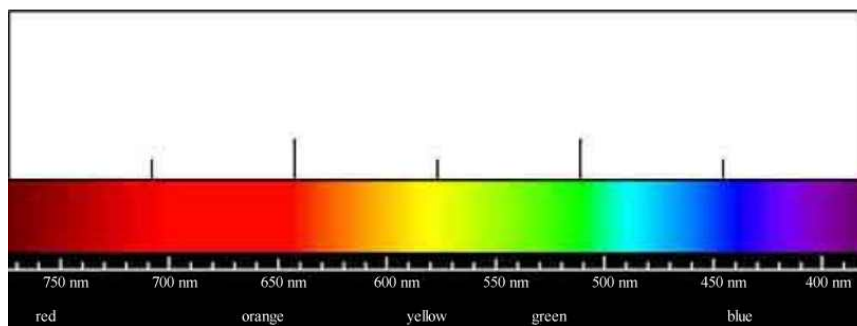
Conclusion:

1. Your TA will inform you which section of this lab you should write up for a conclusion.

Table 1: Wavelengths (in nm) of the emission lines of various elements.

H			He	Ne	Ar	I	Hg	Kr
410.2	H_δ	$n = 6$	453.8	404.4	486.2	404.7	427.4	
434.0	H_γ	$n = 5$	470.4	415.9	511.9	435.8	432	
486.1	H_β	$n = 4$	471.5	416.4	546.5	546.1	436.3	
656.3	H_α	$n = 3$	501.6	482.7	419.1	608.2	577	437.6
			587.6	495.7	434.5	629.4	579.1	445.4
			667.8	503.8	451	656.6		446.4
			706.5	514.5	459.6	695.9		450.2
				534.1	470.2			556.2
				540	518.8			557
				576.4	549.6			587.1
				585.2	565.1			588
				588.2	591.2			599.4
				603	603.2			605.6
				607.4	604.3			642.1
				614.3	605.9			645.6
				616.3	641.6			690.5
				621.7	675.3			758.7
				626.7	696.5			760.1
				640.2				768.5
				650.6				769.4
				659.9				785.5
				692.9				