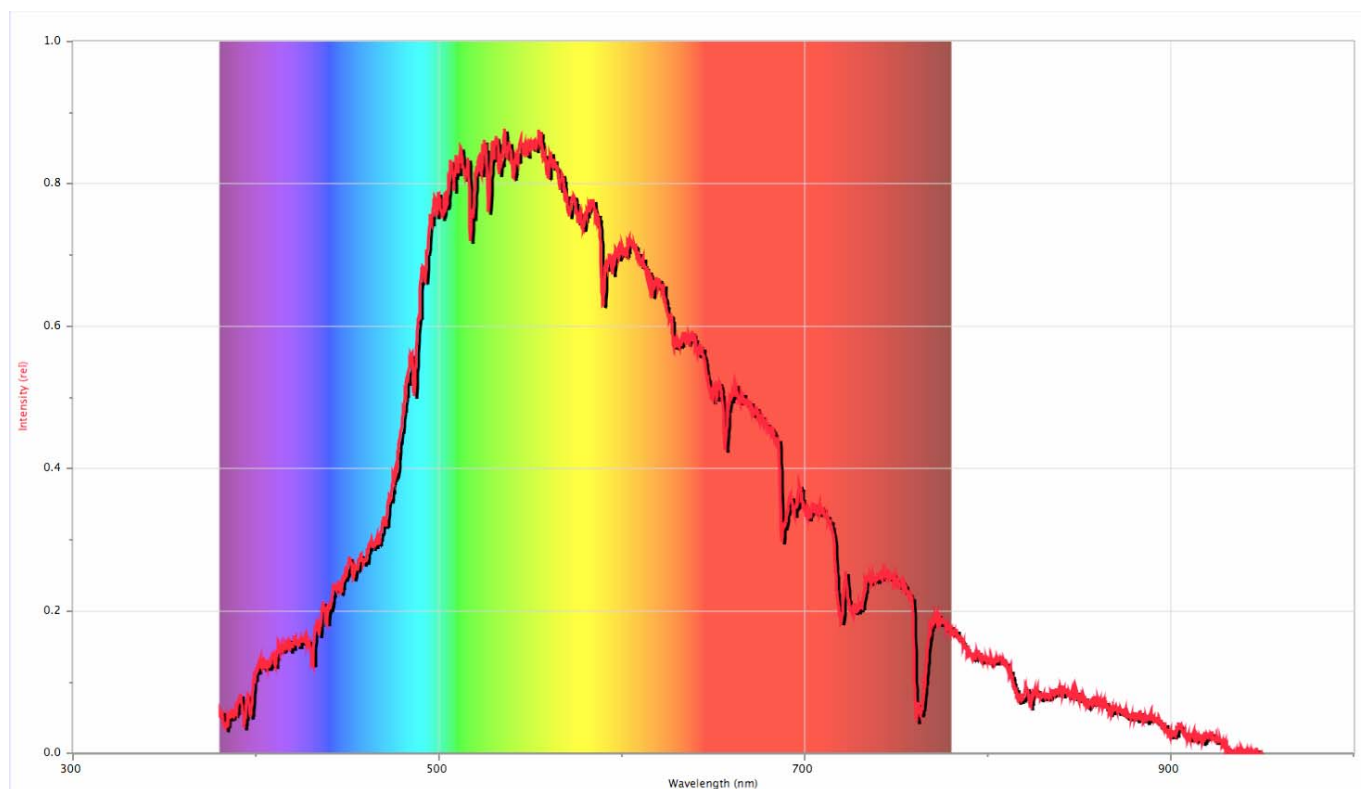# AS 101
Spectroscopy

The Problem

From studying the spectra of stars, we can in principle learn a great deal about their physical properties -- their compositions, temperatures, and densities, the presence of magnetic fields, rotations, and indirectly their masses and ages, can all be determined. In the mid 19th century, Gustav Kirchoff postulated his three ``laws'' of spectroscopy. While Kirchoff's laws consist simply of visual descriptions of the appearance of the spectrum emitted under various conditions, and in many respects they are an oversimplification, they still provide a convenient framework for understanding spectroscopy.

The goal of this exercise is to verify Kirchoff's laws. Under what conditions are the various kinds of spectra produced? How can we use Kirchoff's laws to deduce the properties of stars? Can you identify any elements in the spectrum of the Sun?

Introduction

Most of what we know about the physical nature of stars comes to us from the practice of spectroscopy. Unfortunately, once the light from a faint star has been spread out into a spectrum, it has become so faint that it is difficult to see, unless a large telescope is used. For this exercise, we will examine various types of light sources in the lab to see how the spectrum can change under different conditions of

temperature, composition, and physical state of the material. The types of spectra emitted by materials under various conditions were first described by Gustaff Kirchoff in 1859. He stated his conclusions in the form of three "laws" which still are a useful way to describe the various kinds of spectra.

Kirchoff's First Law: Kirchoff's first law states that any hot substance (i.e. solid, liquid, or gas) under high enough pressure gives off a **continuous spectrum**, that is, it emits photons of all wavelengths. The wavelength (or color) at which the spectrum appears brightest depends on the temperature of the object. You can investigate Kirchoff's first law by using an incandescent bulb connected to a variac (a source of electrical energy whose power level may be varied). The material making up the filament in the bulb provides the continuous spectrum. The variac allows the temperature of the filament to be varied based on the amount of power supplied to the filament.

Kirchoff's Second Law: An **emission line** spectrum is produced by a hot rarified gas, according to Kirchoff's second law. The spectrum given off consists of bright lines at specific wavelengths. These lines are characteristic of the particular substance being used. An emission spectrum can be produced by a discharge tube which contains a small amount of gas and two electrodes. When a high voltage is applied across the electrodes, the gas is heated, and the electrons in the gas are excited into higher energy levels. For an electron to reach a certain energy level it must absorb precisely the right amount of energy for that level. This precise amount of energy is called a quantum of energy. There can be many energy levels for an atom and a different energy levels for different atoms.

In returning to lower energy levels, the gas atoms emit photons of specific energies (therefore specific wavelengths), which can be seen using the spectroscope. Each energy, and therefore wavelength, of light emitted correspond to different drops in energy levels. The energy, **E**, of a photon can be found by multiplying Planck's constant, **h** $= 6.626 \times 10^{-34}$ J·s, by the photons frequency, **f**.

$$E = hf$$
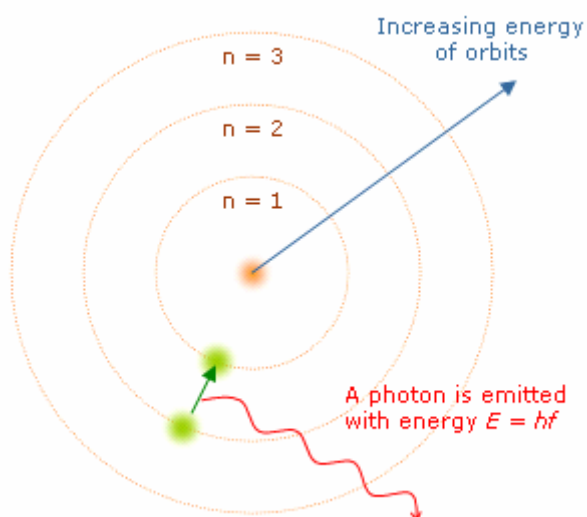
To calculate the wavelength of light emitted by an atom whose electrons have dropped from an excited state to a lower state around the nucleus use:

$$\frac{1}{\lambda_{\text{vac}}} = R_{\text{H}} Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

where

λ_{vac} is the wavelength of the light emitted in vacuum;
R is the Rydberg constant for hydrogen = 10967758.341 m^{-1}
Z is the atomic number, i.e. the number of protons in the atomic nucleus of this element;
n_1 and n_2 are integers such that $n_1 < n_2$.

For an electron falling from the first excited state n = 2 to the ground state n = 1, a photon is emitted with a wavelength, λ = 121.6nm, which corresponds to 1.635 J of energy.

A particularly useful unit when dealing with the small energies of atomic physics is the electron volt, **eV**. An eV is the energy given to an electron by accelerating it through 1 volt of electric potential difference.

$$1 \text{ eV} = 1.6 \times 10^{-19} \text{ J}$$

1.635×10^{-18} J of energy converts to 10.2 eV.

Since each element has different numbers of protons and electrons, each element emits light at characteristic wavelengths. This makes spectroscopy a powerful tool for chemical analysis.

<u>Kirchoff's Third Law:</u> According to Kirchoff's Third Law, an **<u>absorption spectrum</u>** is observed when a cooler gas is situated in front of a hotter material. In this case, the cool gas *absorbs* light emitted by the hotter material, again at specific wavelengths characteristic of the cooler substance. The absorption and emission spectra are complimentary, in the sense that for a specific element, its emission line pattern will exactly match its absorption line pattern. The Sun has an absorption spectrum which can be seen with our spectroscopes.
According to Kirchoff's third law, when a hot, opaque object (or any other source of a continuous spectrum) lies behind a cool, low-density gas, dark absorption lines appear in the otherwise continuous spectrum. This occurs in nearly every star, since light from the interior must pass through the cooler, outer layers. In fact, it was in the spectrum of the sun that William Wollaston first observed dark absorption lines in 1802. Between 1814-1815, Joseph Fraunhofer carefully studied the solar spectrum to find approximately 600 dark lines. Since he was the first to recognize these absorption lines, these solar lines are called "Fraunhofer lines." Some of the "lines" are actually much broader in wavelength than others.

For this lab, you will use a **<u>spectroscope</u>** that uses a **<u>diffraction grating</u>** as the dispersing element to spread the light out into a spectrum of colors. The principle of **<u>diffraction</u>** refers to the spreading or bending of light around an obstacle, such as a slit. Our instruments use gratings with thousands of closely spaced, equidistant, and parallel etched lines on a transparent medium. Incident white light is diffracted by the grating and separated into its constituent wavelengths by angle. Figure 1 represents a traditional spectroscope with a transmission grating showing that the diffracted spectrum is displayed on either side of the central position. In this spectroscope, the light is passed through a slit, the image of which is projected as shown.

1:     Figure 2: Grating Spectroscope
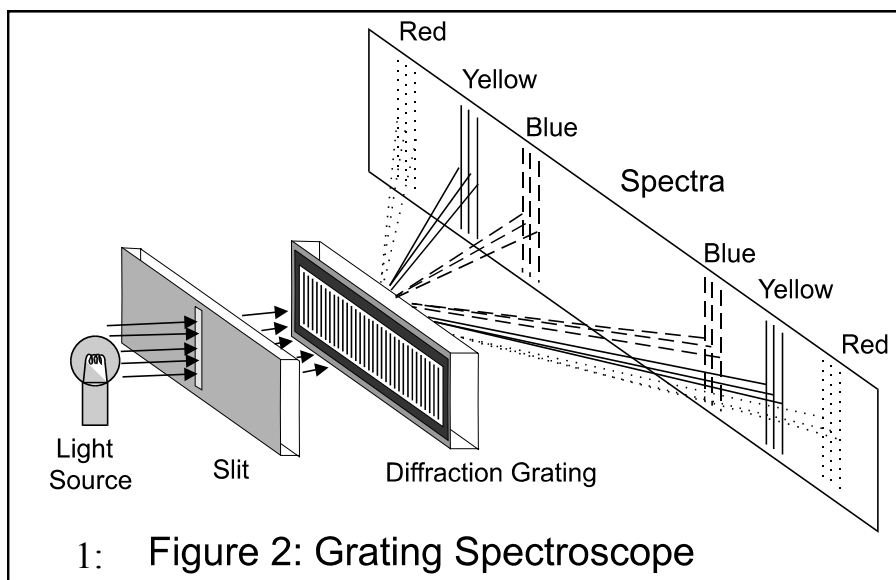
To get a feeling for how a spectroscope works you will use a simple grating spectroscope (Figure 2) that has grating material in a window at the narrow end of the instrument. On the wider end, a small slit is placed to one side. To operate this instrument, look through the window to view the spectrum, while pointing the slit (the side forming a right angle with wide end, with the slit being at the intersection of edges) in the direction of the light source. The spectrum will appear on either side of the slit. There is a scale visible, marked in units of nanometers; the spectrum is seen in projection against the scale.


Figure 2 – Project Star spectroscope

For the main part of the lab we will use a modern, computer-based spectroscope, where the slit is replaced by an optical fiber. After the light is passed through the diffraction grating, it is projected onto

a detector, which sends a signal to a computer. You will use the same "Logger Pro" software you used in the gravity lab to view and analyze the spectrum (see figure 3).

To do this make sure the spectrometer is plugged into the computer via the USB cable. Open *Logger Pro* using the icon on the desktop. Under the *Experiment* tab mouse over to *Change Units* and under *Spectrometer: 1* select *Intensity*.

To observe the spectrum, you need to point the fiber optic at the source. You can take measurements of the spectrum by clicking the *Collect* button (a little green button with an arrow), located above the graph.
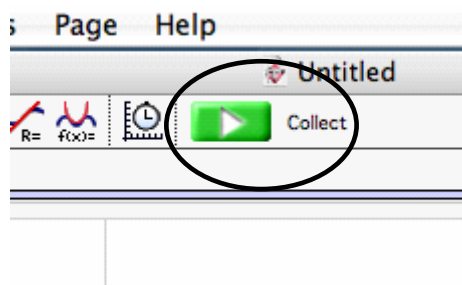
Available Equipment

1. A simple diffraction grating spectroscope.

2. A computer-based grating spectroscope

3. An incandescent light bulb with a rheostat for controlling the voltage to the bulb.

4. Several gas discharge tubes filled with common gases.

5. Sunlight.

Suggested Experimental Procedure:

The goal of this project is to verify Kirchoff's "laws". Since his laws exist primarily as a set of descriptions of the appearance of the spectrum under various conditions, this exercise will consist mainly of a set of demonstrations of the phenomena associated with his laws. So there will be three separate parts to the lab, demonstrating each of his laws separately.

Continuous spectrum

A variac (which adjusts the voltage flowing through an electrical circuit) will be set up in series with a light bulb. As the voltage on the variac is changed, the increased current flowing through the bulb heats the filament so that it glows brighter and bluer.

Try several voltages and examine the bulb through the spectroscope. Describe the changes you observe. In particular, do you see any changes in the relative intensity of the red end of the spectrum vs the blue end? What other changes (if any) do you see, as the voltage is changed? Can you think of any quantitative ways to test the nature of the continuous spectrum?

Emission spectrum

Several gas discharge tubes will be set up in the lab. These tubes contain various gases, as marked on the tube. When an electric current is passed through the tube, the atoms of the gas are excited, and they emit light at the characteristic wavelengths for that gas. Examine some of these tubes with your spectroscope. Sketch the spectrum of a few of them on the charts at the end of this lab writeup. Record the wavelengths of some of the most prominent spectral lines in each spectrum, using the scale visible through the spectroscope. Pay particular attention to the sodium and hydrogen lamps

Absorption spectrum

The easiest way to see an absorption spectrum is to look at the spectrum of the Sun. Even on a cloudy day, the daylight has the spectrum of the Sun - simply aim the spectroscope out a window and record the solar spectrum in the computer. Some of the stronger absorption lines will be visible on your graph. Record the wavelengths of the most obvious absorption lines. Do any of these lines match the wavelengths you found for the hydrogen lines or the yellow sodium line you saw in the lab?

You will also be able to observe absorption spectra by shining the incandescent light through several fluids in the lab. The fluids contain complex molecules that absorb light over fairly broad wavelengths. Record the spectra of a few of these.

If the Sun is out, you will be able to observe the sodium region of the spectrum with the solar spectrograph. Compare what you see there with what you found earlier. You may also want to examine the solar spectrum on the wall of the lab. This spectrum was made with our solar spectrograph by an undergraduate astronomy student some years ago. It is a beautiful piece of work and shows the spectrum of the Sun in great detail.
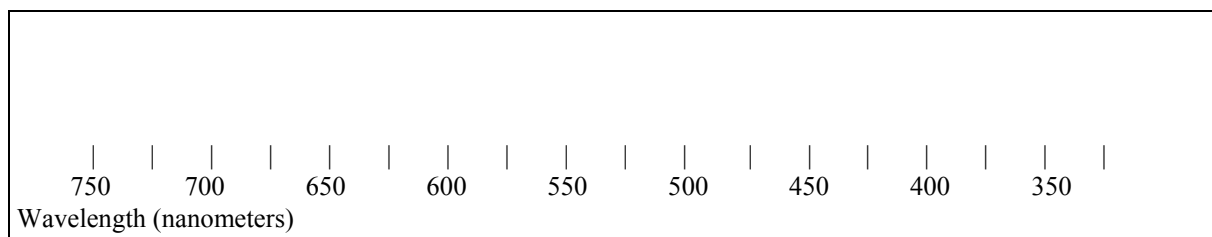
Analysis and discussion

You should begin your analysis by considering your identification of solar spectral lines. With the help of the solar spectrum on the lab wall and other sources, you should be able to identify most of the spectral lines you were able to record. However, there is a real possibility that the scale on your spectroscope could be out of alignment. Use the sodium lamp or other sources to calibrate the scale of your spectroscope – In this case the calibration should consist of a constant to add to (or subtract from) your measured wavelengths to get the standard wavelengths. For reference, the two sodium "D" lines have wavelengths of 589 and 596 nanometers. Keep in mind that with your spectroscope, you probably do not see two lines, but simply a blend of the two lines. Estimate how accurately you can measure wavelengths with your spectroscope.

Overall, this lab is rather less quantitative than most of the labs. Your report should describe in qualitative terms how your observations support (or do not support) Kirchoff's concepts of spectroscopy. Be sure to discuss the details of the procedure you actually followed
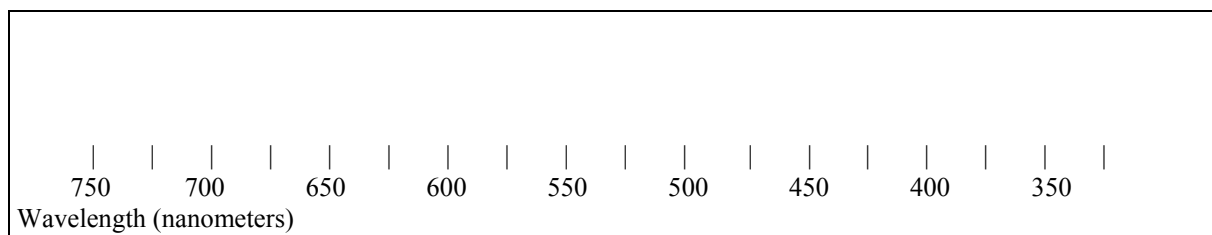
Records and data

Your "data" for this lab will consist primarily of notes you took while observing the various spectra. You will also be able to get printouts of some of the spectra. Where appropriate, you might need to include your own diagrams. For the hydrogen and sodium tubes you should make a table of wavelengths and the relative intensities of the various lines.
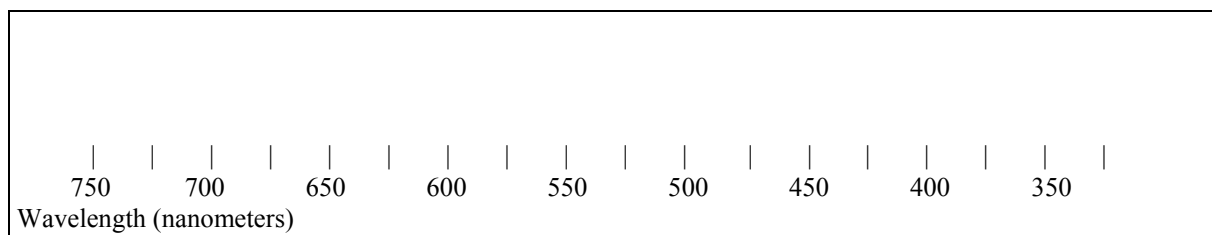
In addition to your notes you should record the positions of the most prominent spectral lines in the diagrams on the following pages. (If you wish, there will be colored pencils available for your diagrams, although they are not necessary):

Spectral Line Diagrams - Page 1

| | | | | | | | | | | | | | | | | | | | | |
750   700   650   600   550   500   450   400   350
Wavelength (nanometers)

Description of Source_____

| | | | | | | | | | | | | | | | | | | | | |
750   700   650   600   550   500   450   400   350
Wavelength (nanometers)

Description of Source_____

| | | | | | | | | | | | | | | | | | | | | |
750   700   650   600   550   500   450   400   350
Wavelength (nanometers)

Description of Source_____

Spectral Line Diagrams – Page 2

| | | | | | | | | | | | | | | | |
| 750 | 700 | 650 | 600 | 550 | 500 | 450 | 400 | 350 |
Wavelength (nanometers)

Description of Source_____

| | | | | | | | | | | | | | | | |
| 750 | 700 | 650 | 600 | 550 | 500 | 450 | 400 | 350 |
Wavelength (nanometers)

Description of Source_____

| | | | | | | | | | | | | | | | |
| 750 | 700 | 650 | 600 | 550 | 500 | 450 | 400 | 350 |
Wavelength (nanometers)
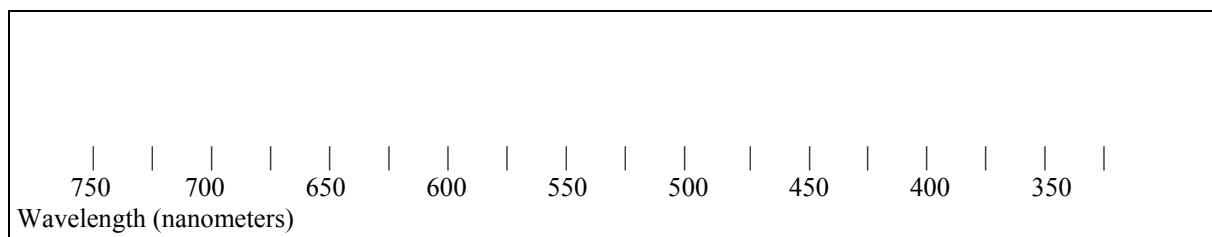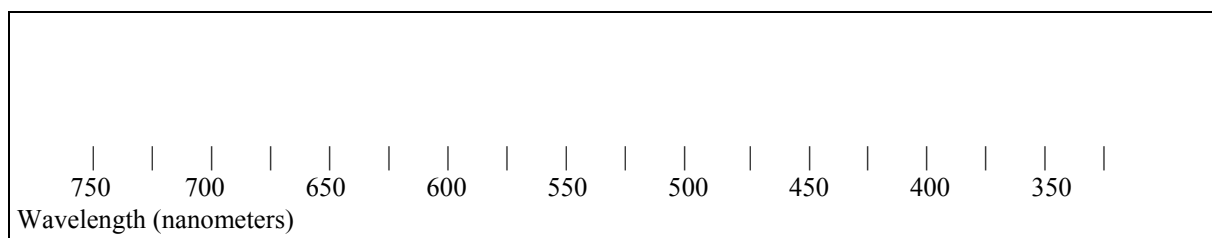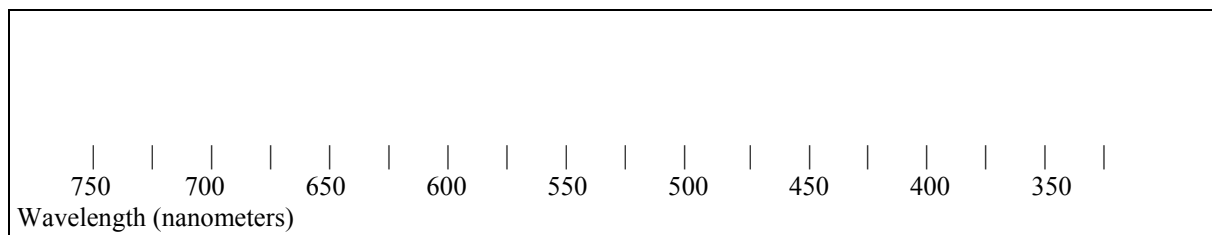
Description of Source_____

Fill out the chart for electrons falling from the different energy levels of Hydrogen. Were you able to see these emission lines?

$n = 7$

$n = 6$

$n = 5$

$\lambda = \underline{\hspace{2cm}}$ $E = \underline{\hspace{2cm}}$

$n = 4$

$\lambda = \underline{\hspace{2cm}}$ $E = \underline{\hspace{2cm}}$

$n = 3$

$\lambda = \underline{\hspace{2cm}}$ $E = \underline{\hspace{2cm}}$

$n = 2$