SPECTROSCOPY

"I ask you to look both ways. For the road to a knowledge of the starts leads through the atom; and important knowledge of the atom has been reached through the stars." - A.S. Eddington

SYNOPSIS: In this lab you will explore different types of emission spectra and use a spectroscope to identify an unknown element. You will also observe and analyze the spectrum of the Sun using the observatory's heliostat to study the Sun's composition.

EQUIPMENT: Hand-held spectroscope; spectrum tube power supply and stand; helium, neon, nitrogen, air, and "unknown" spectrum tubes; incandescent lamp; and the heliostat.

WARNING: There is high voltage on the spectrum tube. You can get a nasty shock if you touch the ends of the tube while it is on. Also, the tubes get hot and you can burn your fingers trying to change tubes. Let the tubes cool or use paper towels to handle them.

Most of what astronomers know about stars, galaxies, nebulae, and planetary atmospheres comes from **spectroscopy**, the study of the colors of light emitted by such objects. Spectroscopy is used to identify compositions, temperatures, velocities, pressures, and magnetic fields.

An atom consists of a nucleus and electrons which "orbit" around the nucleus. An atom emits energy when an electron jumps from a high-energy orbit around the nucleus to a smaller low-energy orbit.   The energy appears as a **photon** of light having an energy exactly equal to the *difference* in the energies of the two electron levels. A photon is a wave of electromagnetic radiation whose wavelength (distance from one wave crest to the next) is inversely proportional to its energy: high-energy photons have short wavelengths, low-energy photons have long wavelengths. Since each element has a different electron structure, and therefore different electron orbits, *each element emits a unique set of spectral lines*.
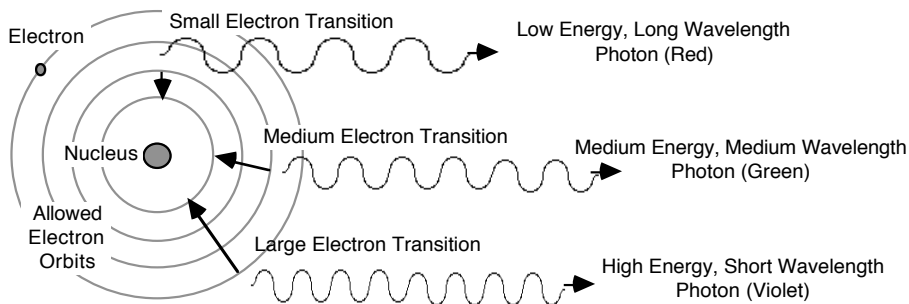
The light produced when electrons jump from a high-energy orbit to a smaller low-energy orbit is known as an **emission line**. An **absorption line** is produced by the opposite phenomenon: a passing photon of light is absorbed by an atom and provides the energy for an electron to jump from a low-energy orbit to a high-energy orbit. The energy of the photon must be exactly equal to the *difference* in the energies of the two levels; if the energy is too much (wavelength too short) or too little (wavelength too long), the atom can't absorb it and the photon will pass by the atom unaffected. Thus, the atom can selectively absorb light from a continuum source, but only at the very same wavelengths (energies) as it is capable of emitting light. An absorption spectrum appears as dark lines superimposed against a bright continuum of light, indicating that at certain wavelengths the photons are being absorbed by overlying atoms and do not survive passage through the rarefied gas.

The human eye perceives different energies (wavelengths) of visible light as different *colors*. The highest energy (shortest wavelength) photon detectable by the human eye has a wavelength of about 4000 Angstroms (one Angstrom equals 10^{-10} meters) and is perceived as "violet". The lowest energy (longest wavelength) photon the eye can detect has a wavelength of about 7000 Angstroms, and appears as "red".

A **spectroscope** is a device that allows you to view a spectrum. Light enters the spectroscope through a slit and strikes a grating (or prism) which disperses the light into its component colors (wavelengths, energies). Each color forms its own separate image of the opening; a slit is used to produce narrow images, so that adjacent colors do not overlap each other.
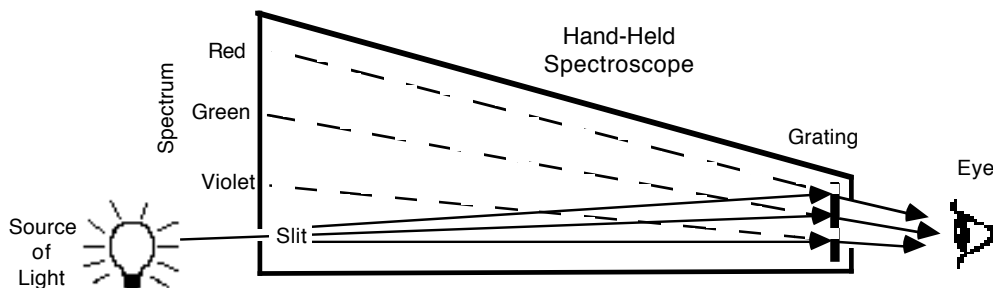
A **grating** is a sheet of material with thousands of evenly spaced parallel openings. In general, the crests and troughs of the light waves passing through adjacent openings *destructively interfere* with each other: they cancel each other (crest on trough) so that no wave survives. At a unique angle for each wavelength, however, the waves *constructively interfere* (crests add to crests, troughs to troughs); at that angle we see a color image formed by that wavelength of light. Longer wavelengths of light are "bent" (**diffracted**) the most, so a spectrum is formed with violet seen at small angles, and red at large angles.

Part I. Continuum and Emission Line Spectra

I.1   Install the *helium* discharge tube in the power supply and turn it on. Describe the apparent color of the glowing helium gas. Now view the gas using the hand-held spectroscope (*held horizontally so that the vertical slit is aligned with the vertical glowing tube*), and note the distinctly separate colors (**spectral lines**) of light emitted by the helium atom. Make a sketch (including wavelengths and units!) of the spectrum and label the colors. Compare the apparent overall color of the glowing gas with the collection of actual colors contained in its spectrum.

I.2   Turn off the helium tube and replace with the *neon* gas tube. Judging from the appearance of glowing neon, what wavelengths would you expect to dominate its spectrum? Observe and sketch the spectrum using the spectroscope as before. Judging from the number of visible energy-level transitions (lines) in the neon gas, which element would you conclude has the more complex atomic structure: helium or neon?

A number of neon spectral lines are so close together that they are difficult to separate or *resolve*. This blending of lines becomes even more dramatic in *molecules* of gas: the constituent atoms interact with each other to increase the number of possible electron states, which allows more energy transitions, which in turn produces wide molecular **emission bands** of color.

I.3   Install the *nitrogen* gas tube containing diatomic (two-atom) nitrogen (N_2) molecules. Observe and sketch the spectrum, and identify the broad emission bands caused by its molecular state.

A *solid* glowing object such as the tungsten filament of an incandescent lamp will not show a characteristic atomic spectrum, since the atoms are not free to act independently of each other. Instead, solid objects produce a **continuum spectrum** of light regardless of composition; that is, *all* wavelengths of light are emitted rather than certain specific colors.

I.4   Look through the spectroscope at the *incandescent lamp* provided by your instructor. Confirm that no discrete spectral lines are present. What color in the spectrum looks the brightest? How does this color compare to the overall color of the lamp (i.e. the lamp's color without using the spectroscope)?

Fluorescent lamps operate by passing electric current through a gas in the tube, which glows with its characteristic spectrum. A portion of that light is then absorbed by the solid material lining the tube, causing the solid to glow or **fluoresce** in turn.

I.5   Point your spectroscope at the ceiling *fluorescent* lights, and sketch the fluorescent lamp spectrum. Identify which components of the spectrum originate from the gas, and which from the solid.

Part II. Identifying an Unknown Gas

II.1   Select one of the unmarked tubes of gas (it will be either hydrogen, mercury, or krypton). Install your "mystery gas" in the holder and inspect the spectrum. Make a sketch of the spectrum and label the colors.

II.2   Identify the composition of the gas in the tube by comparing your spectrum to the spectra described in the table below. (Strongest lines are shown in **boldface** type. The numbers to the left of each color are the wavelengths of the spectral lines given in Angstroms).

Hydrogen	
6563	**Red**
4861	**Blue-Green**
4340	Violet
4101	Deep Violet (dim)

Mercury	
6073	Orange
5780	**Yellow**
5461	**Green**
4916	Blue-Green (dim)
4358	**Violet**
4047	Violet (dim)

Krypton	
6458	**Red**
5871	**Yellow-Orange**
5570	**Green**
4502	Violet (dim)
4459	Violet (dim)
4368	Violet (dim)
4320	Violet (dim)
4274	Violet (dim)

II.3   Install the tube of gas marked "*air*" and inspect the spectrum. Compare it to the spectrum for nitrogen. What does this tell you about the composition of our atmosphere? What molecule(s) do you think are responsible for the spectral lines in air?

Part III. The Solar Continuum

Now you'll apply what you learned about gases in the laboratory to gases in the Sun.   First, let's study the Sun's continuum spectrum. The gas in the interior of the Sun is so dense that the atoms cannot act independently of each other and therefore they radiate a continuum spectrum just like a solid. The solar continuum is a **blackbody spectrum**. Like all blackbody spectra, the shape of the solar continuum depends only on its temperature.

The graphs below show the blackbody spectra seen for *any* object (not just stars!) with different temperatures. The temperatures for the graph on the left are chosen to show the large variation in radiated energy for minor variations in temperature. The graph on the right uses a logarithmic scale to encompass a larger range, and shows the temperatures for "O" stars (50,000 K - the hottest known stars), the Sun (5783 K), and a person (310 K).
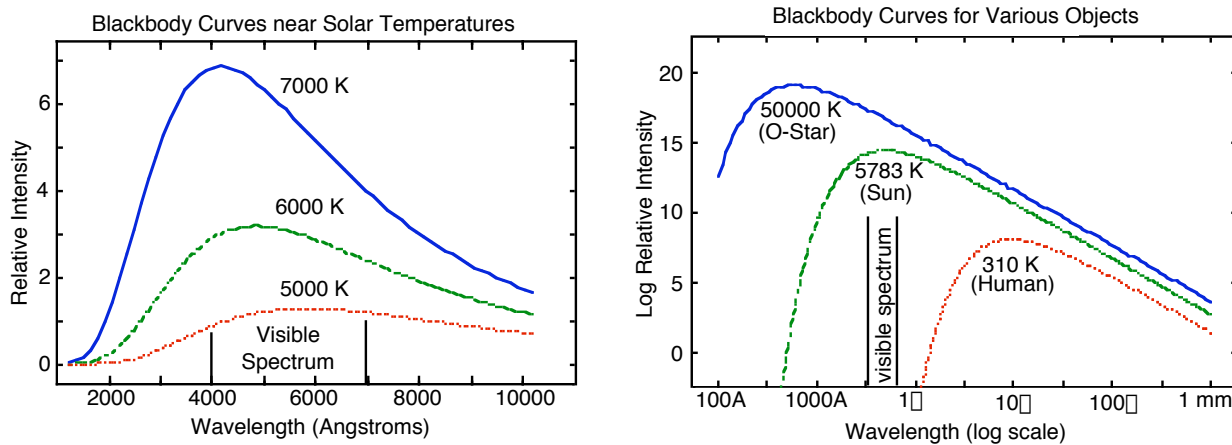
III.1   Why do you suppose we humans have evolved to see light in the wavelength range that we do (4000 to 7000 Angstroms?) If the Sun were hotter, would it be useful for us to see at longer or shorter wavelengths? Why can't we see each other "glowing in the dark"?

An object's temperature can be determined from the wavelength at which the peak energy occurs using **Wien's Law**:

$$\text{Temperature} \ = \frac{2.90 \times 10^{7} \ (\text{degrees kelvin}_{x} \ \text{Angstroms})}{\text{Peak Wavelength in Angstroms}}$$

By measuring the peak energy of the solar continuum, we can determine the temperature of the surface of the Sun.

III.2   Where in the solar spectrum does the continuum emission *appear* (with your eye) the brightest? Estimate the wavelength using the Fraunhofer Solar Spectrum Chart below the spectrograph screen.

III.3   Using a photometer (lightmeter) provided by your instructor, measure where the continuum is the brightest. Estimate the wavelength. How does this compare to your estimation with your eye?

III.4   Using Wien's law, calculate the temperature of the Sun using your measured estimate for the peak wavelength. How does this compare to the accepted value of 5783 K?

Part IV. Solar Absorption Lines

In the Sun we see *both* a continuum spectrum and absorption lines from gases in the tenuous solar atmosphere, which selectively absorb certain photons of light where the energy is "just right" to excite an atom's electron to a higher energy state.

Of course, an atom can't absorb light if the atom isn't there. Thus, we can identify the elements that make up the Sun simply by comparing the absorption lines in the solar spectrum with emission lines from gases here on Earth. Your lab instructor will dim the solar spectrum by narrowing the entrance slit, and switch on comparison lamps containing *hydrogen* and *neon*; their emission spectra appear above and below the solar spectrum. If an element is present in the Sun (and if the temperature of the gas is "just right"), its absorption line will appear at the same wavelength as the corresponding emission from the comparison source.

IV.1   Which comparison spectrum (top or bottom) is due to hydrogen, and which is neon? What is your evidence?

IV.2   Do you see evidence that hydrogen gas is present in the Sun? What about neon?

Even though your eye can't see it, the solar spectrum extends beyond the visible red portion of the spectrum to include **infrared** light, and beyond the violet to include **ultraviolet** light.   The spectrum of ultraviolet light can be made visible by *fluorescence* (remember the solid material in the overhead tube lights?).

IV.3   Hold a piece of bleached white paper up to the viewing screen where ultraviolet light should be present in the spectrum. Note that there are *many* additional absorption lines in the solar spectrum at wavelengths that are not normally visible to us! Use the Solar Spectrum Chart to identify the element responsible for the two extremely broad, dark lines in the ultraviolet.

IV.4   Using the Fraunhofer chart, identify each of the following lines in the solar spectrum - for each, give its wavelength, color, and include a brief descriptive phrase (dark, broad, narrow, fuzzy, etc.).

Line Name and Element	Abundance per million	Wavelength (angstroms)	Color	Descriptive Appearance	Reason
B Oxygen (O_2) - molecular	none				
C Hydrogen alpha (H α)	900,000				
D Sodium (Na)	2				
E Iron (Fe)	40				
F Hydrogen beta (H β)	900,000				
H Calcium (Ca II) - ionized	2				
K Calcium (Ca II) - ionized	2				

It *seems* reasonable to assume that stronger (darker, wider) absorption lines would indicate that there are more atoms of that element present in the Sun - but as shown in the "abundance" column of the table, we would be misleading ourselves. Although an atom must be *present* in order to show its "signature", that doesn't necessarily mean that the element is *abundant* - otherwise, we would probably assume (incorrectly) that the Sun is mostly made out of calcium!

The analysis of spectral features is an extremely complex task, and there are several important factors to consider. First of all, even if *many* atoms of an element are present, the temperature may be too *cool* to "excite" them, so that they won't emit or absorb (remember, the gas in the discharge tubes was completely invisible until you turned on the power). Or, if the temperature is too *hot*, the atoms will be ionized (electrons stripped away) so that the electrons responsible for the absorption won't be part of the atom anymore. Or, it could be that a spectral feature that we see didn't really come from the Sun at all, but from the gasses in our own Earth's atmosphere!

 Through careful analysis, physicists have figured out the temperature ranges necessary for certain elemental species to "show themselves" through absorption, as indicated below.
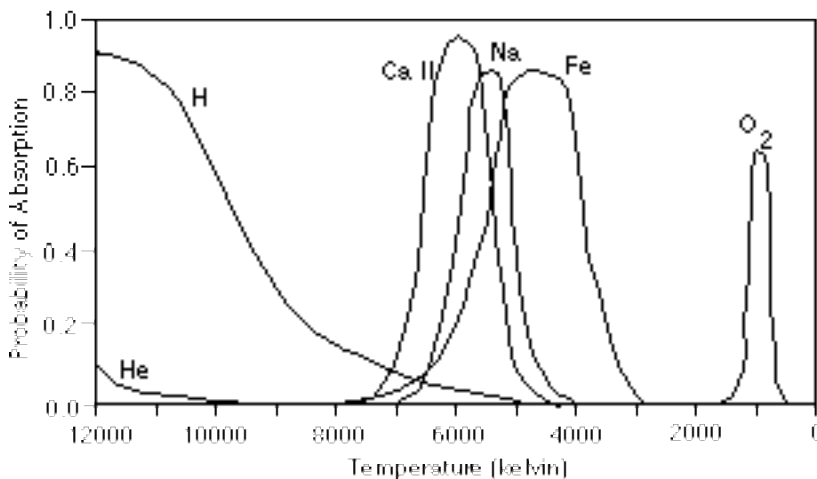
IV.5   For each spectral line in the table on the previous page give a reason (A, B, C, or D below) why you think it appears in the solar spectrum.

   A. Although only a very small fraction of the atoms absorb light at any one time, this atom is very abundant and therefore shows its signature.

   B. This spectral line is actually several lines that are closely spaced together. The sum of the lines makes it appear to be stronger than it really is. It's not really abundant in the Sun.

   C. This spectral line is strong because the Sun's temperature is "just right" for this element to absorb light at this wavelength, even though there aren't that may atoms there.

   D. The spectral "line" is actually an absorption band from molecules in Earth's atmosphere rather than in the solar atmosphere. The Sun is too hot for molecules to exist.

IV.6   Although 10% (100,000 atoms out of a million) of the Sun is composed of helium (He), we don't see *any* helium lines in the visible solar spectrum. Inspect the graph of absorption probabilities, and explain why not.

IV.7   However, during total solar eclipses astronomers *do* see helium in the solar **corona**, the extremely tenuous outermost region of the Sun's atmosphere (in fact, the element was named for the Greek word *helios*, meaning "the Sun". What does this imply about the temperature of the gas in the corona? (If the answer surprises you, you're not alone - it also surprised astronomers!)