Lab 2. The Nature of Light

1 Introduction

Astronomy is rather different from most other physical sciences. In physics or chemistry, for example, in order to discover some property of a material or some law governing the behavior of matter, one only has to build and perform the appropriate experiment and analyze the results. In astronomy one cannot perform such experiments; one can only observe experiments that the universe has already set up for us. Thus, an astronomer must not only be able to analyze the results of an experiment, but he or she must also work backwards to figure out what the experiment actually is!

Astronomers also differ from scientists in other fields in that they have only one sense. A chemist may be able to smell a chemical, or feel a precipitate, or hear an explosion, but for an astronomer, these senses do not exist. Except perhaps for solar system measurements (where we can send astronauts to go by rocket ship), all the astronomer can ever do with the universe is look at it. This being the case, astronomers have to understand more than a bit about light.

In this lab, you will perform your first measurement of light, and learn one of the ways that astronomers use light to learn about the universe. But first, we will present some background as to how light actually works.

2 What light is

Light is unlike anything else in the universe, and that makes it a bit difficult to describe. Normally, we (or at least I) learn things by analogy, *i.e.,* by comparing some new experience to something we are already familiar with. However, there is nothing in our experience like light. We can try a few analogies, and many times these analogies will work, but no one analogy, or model, works perfectly for light. Keep this in mind as we proceed.

Light can sometimes be thought of as a particle, or *photon,* carrying a "packet" of energy. However, in many cases, the best way to think of light is as a wave. Like a water wave at the beach, a light wave can be described as having two characteristics: the distance between two waves, or *wavelength* (abbreviated by λ, the Greek letter lambda), and the wave frequency, or number of waves passing you each second (abbreviated by ν, the Greek letter nu). Fig. **??** shows how these quantities are defined.

Wavelength and frequency are related in a fairly obvious manner. In the case of water waves, if two waves are 5 meters apart and 10 waves pass you each minute, then the water must be traveling

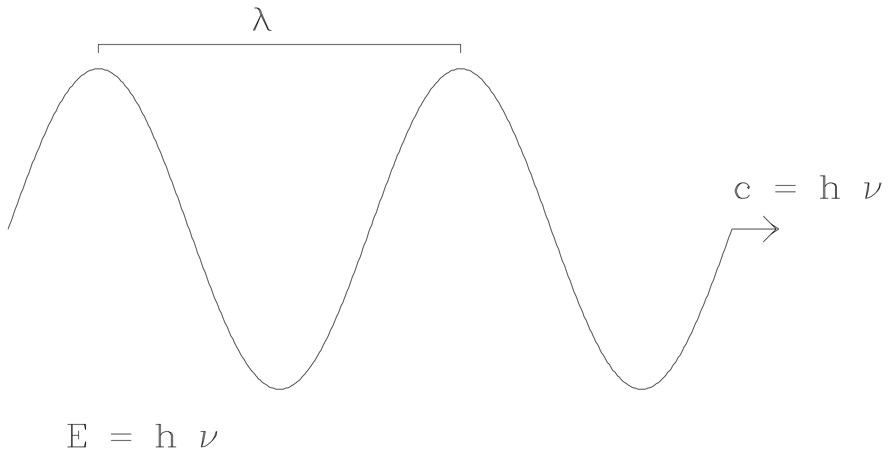

Figure 1: Basic properties of a light wave.

at $5 \times 10 = 50$ meters/minute. In other words, the speed of the wave,

$$c = \lambda\nu \tag{1}$$

This is true for any wave, but for light waves, there is an additional property. The speed of light is ALWAYS the same, and the letter c is nearly always used to denote its value: $c = 3 \times 10^8$ m-sec^{-1}. Thus, the wavelength of light uniquely defines its frequency, and vice-versa. The higher the frequency, the smaller the wavelength. Furthermore, there is also a unique relationship between the wavelength (or frequency) of light, and how much energy that light carries. This formula is simply

$$E = h\nu = hc/\lambda \tag{2}$$

where h is a number called the *Planck constant* and is 6.63×10^{-34} joule-sec. Thus, the longer the wavelength, the less energy the light carries.

The wavelength of light can be anything, from several hundred meters long, to the size of an atomic nucleus. Our bodies don't react to extremely long wavelength light, but this type of light can move electrons in wires, and such waves are usually called *radio* waves. Light of shorter wavelengths, say a few microns (μ), does react with our bodies, and when subjected to this type of light, our nerves will send our brain the message, "heat." Light of much shorter wavelength ($< 0.3\mu$) may penetrate our skin, and damage our cells. Depending on how short the wavelength is, we call this ultraviolet, x-ray, or γ-ray (gamma-ray) light. Between all this, however, is a very small wavelength region where light reacts with our eyes. We call light with a wavelength between $0.3\mu < \lambda < 0.7\mu$ optical light. In this region are the colors of the rainbow: violet and blue at the short wavelength end, and red light at the long wavelength end. The entire *electromagnetic spectrum* with an enhanced look at the optical spectrum is presented in Fig. **??**.

Radio astronomers usually use centimeters as their basic unit of wavelength, infrared astronomers give wavelengths in terms of microns (10^{-6} meters), and physicists and engineers usually present their measurements in nanometers (10^{-9} meters). However, for historical reasons, optical astronomers have always quoted their wavelengths in a unit call the Angstrom. One Angstrom, abbreviated Å, equals 10^{-10} meters. Blue light is about 4000 Å; red light is around 7000 Å.

3 Making light via the Blackbody Law

Perhaps the best way to consider what light is and how it is made is to recall Newton's first law of motion: a body in motion tends to stay in motion, or a body at rest tends to stay at rest, unless acted upon by an outside force. This statement, which was created to describe the motions of bodies on earth and the planets of the solar system, also applies to elementary particles, such as electrons and protons, and it is at this level that light is created.

Consider a free electron, and a free proton. When an electron, which has a negative charge, flies by the proton, its path will change as a result of the positive attraction. The electron will "fall" a bit and give up some of its energy. When it does this, it emits a photon. The wavelength of this photon will depend on how far it falls. Now, consider another free electron, flying past a proton. Perhaps the photon that was emitted by the first electron will strike the second electron. This electron will absorb the energy, move faster, and therefore move farther away from the proton. (How fast or far it moves, of course, depends on the energy, or wavelength, of the photon.) Now consider a whole lot of free electrons moving around. Once again, their motions will continually be deflected by the attraction of protons or (more likely) the repulsion of other flying electrons. With each deflection, a photon will be absorbed or emitted, and again, the energy of the photon will depend on the amount of deflection

Obviously, in the case of an ensemble of free electrons, photons will be emitted with all sorts of wavelengths. Just as obviously, what these wavelengths are, and how much total light is emitted will depend on the average motion of the electrons. If the electrons are moving slowly, then most of the deflections will be slight and there will be a relatively long time between collisions; thus, what little light is emitted will be at long wavelengths. If the kinetic energy of the electrons is high, however, there will be many deflections, and some of the deflections will be violent. A lot of short wavelength photons will result. Thus, the spectrum of light emitted depends on the velocity of the electrons, or, more properly, on their temperature. The higher the temperature, the faster the electrons move and the more short wavelength photons are emitted.

There is an equation, of course, which describes the spectrum as a function of temperature. Its derivation comes from quantum mechanics, and is beyond the level of this course (about 2 years beyond). Nevertheless, we can use the equation, since it describes the amount of light emitted by

Figure 2: The main regions of the electromagnetic spectrum.

Figure 3: A graph showing the basic shape of the Planck function.

any body, at any temperature, regardless of size, shape, or composition of the material. If a body is at temperature, T, then the amount of energy each square meter of its surface emits each second, as a function of wavelength, is

$$B_\lambda(T) = \frac{2hc^2}{\lambda^5}\frac{1}{e^{hc/\lambda kT} - 1} \tag{3}$$

where c is the speed of light, h is the Planck constant, and k is Boltzmann's constant, 1.38×10^{-23} joules °K^{-1}. Equation **??** is called the Planck function, or the blackbody law. A plot of the function is displayed in Fig. **??**.

Obviously, the form of the Planck function is complicated (although it is not difficult to program into a computer). The implications of the law, however, are much simplier to express. For example, how does the peak of the Planck function change with temperature? For this, we can just take the derivative of the function and set it equal to zero. (I leave this as a mathematical exercise to the motivated reader.) If you do this, and plug in the numbers, you find that

$$\lambda_{\mathrm{max}} = 0.2897/T \tag{4}$$

where λ_{max} is expressed in centimeters. This is just what we concluded above: the higher the temperature, the more short wavelength light is emitted. Equation ?? is sometimes called Wien's law.

Question 1 *Human beings, like everything else, emit blackbody radiation. In what region of the spectrum do we radiate most of our light?*

Question 2 *The Sun has an effective temperature of about 6000° K. At what wavelength (or color) does it emit most of its light?*

Another useful quantity to know is how much total energy is emitted by a body at all wavelengths. This is just the integral under the black body curve

$$B(T) = \int_0^\infty B_\lambda(T)d\lambda = 2hc^2 \int_0^\infty \frac{1}{\lambda^5} \frac{1}{e^{hc/\lambda kT} - 1} d\lambda = \frac{2k^4 T^4 \pi^5}{15h^3c^2} \tag{5}$$

or

$$B(T) = \sigma T^4, \tag{6}$$

where σ is called the *Stefan-Boltzmann constant*, and is equal to 5.67×10^{-8} watts-m^{-2} K^{-4}. Equation ?? is the Stefan-Boltzmann law, and it simply says that hotter bodies emit much more light than do cool bodies.

Equations ?? and ?? give us a way to probe the temperatures and sizes of stars. By finding the peak of a star's spectrum, $\lambda_{\rm max}$, we can use Wien's law to estimate its temperature. Moreover, the Stefan-Boltzmann law tells us exactly how much energy each centimeter of that star is emitting, so we can use the total luminosity of the star to measure the surface area that is actually radiating. In other words, for spherical stars (and we'll assume all stars are spherical)

$$L = 4\pi R^2 \sigma T^4 \tag{7}$$

Question 3 *Estimate how much energy a human being emits each second.*

4 Making an emission line spectrum

If light were only created *via* the blackbody law, then we could tell the sizes and temperatures of the emitting objects, but not what they were made of. A 1-meter sphere of iron at 3000 K emits the same *continuous blackbody spectrum* as a 1-meter sphere of lead at 3000 K. Fortunately, nature has provided an additional mechanism for making light. To understand it, we must examine the structure of the atom.

When an electron is unattached to any atom, it is free to emit and absorb any photons it pleases. However, this is not true if the electron is part of an atom. In the Bohr model of the atom (named after the famous physicist Niels Bohr, developer of this model), electron orbits are *quantized;* that is, they can only exist at discrete energy levels.

Figure 4: The Bohr model of the hydrogen atom, the simplest atom. A single electron orbits the nucleus (with a proton) in orbits $n = 1, 2, 3$ and so forth.

Fig. **??** displays the Bohr model for the first few levels of hydrogen. In the center of hydrogen is a proton, which exerts an attractive force on its electron. The potential energy of the system is

$$E = -\frac{q_1 q_2}{r_1} \tag{8}$$

where q_1 and q_2 represent the charge of the electron and proton, and r_1 is the radius of the orbital. Suppose something (say, another electron) collides with this electron. When this happens, some of the energy of the collision will be given to the bound electron, and this electron will be pushed to a higher energy level. After a short time (maybe 10^{-7} seconds), the electron will drop back down to the lower level (this is sometimes termed "decay"), and give up its excess energy in the form of a photon. However, since the distances between the energy levels of an atom are discrete, the amount of energy in this photon is also discrete. In other words

$$\Delta E = \frac{q_1 q_2}{r_1} - \frac{q_1 q_2}{r_2} \tag{9}$$

Now consider an ensemble of these atoms, all undergoing collisions, and all emitting photons when their electrons decay. The spectrum which results from this process will not be a continuous spectrum. Instead all the emitted photons will have energies which correspond exactly to the differences between the Bohr atom levels. An *emission* spectrum will be produced.

Real atoms have hundreds (if not thousands) of energy levels, but the principle described above still exists: when electrons are bound to atoms and decay to lower levels, they are constrained to release photons that correspond exactly to the energy differences between the levels. Thus, by observing a series of *emission lines* from an ensemble of atoms, you can measure the energy levels of those atoms. More importantly, since *no two elements have the same set of energy levels* the emission lines you see tell you what type of atoms are doing the emitting. Thus, you can tell the difference between a tube of hydrogen and a tube of argon by running an electric current through the tube and comparing the different wavelengths which are produced.

In astronomy, emission spectra are most often seen from gas surrounding extremely hot stars. High energy photons from the stars *ionize* the gas by providing enough energy to kick electrons completely out of their atoms. When these electrons recombine into their atoms, they cascade down the energy levels and produce a spectrum with many bright emission lines.

5 Making an absorption line spectrum

In the above sections we have talked about the continuous spectrum made by a blackbody and the emission spectrum made by decaying electrons in an atom. There is one other important type of spectrum to be considered, which is really a combination of the two. Again, consider an ensemble of Bohr atoms, but this time place them between you and a continuous blackbody source. Although photons of all wavelengths (from the continuum source) pass the atoms, only those photons with energies that precisely corresponds to the difference between atomic energy levels can be absorbed. Electrons cannot absorb half the energy of a photon — it's either all or nothing. If a photon has too little energy, or too much energy, the electron *must* ignore it. The result is an *absorption line spectrum,* where you see all the light of a blackbody spectrum *minus* the absorbed photons. The missing wavelengths correspond exactly to the energy levels of the absorbing atoms, and by noting the wavelengths, you can determine what element is doing the absorbing.

In astronomy, absorption lines are usually seen in the spectra of stars. Although the interiors of stars emit as blackbodies, stellar atmospheres are gaseous and relatively cool (at least compared to the inside of the star). This, while passing through the atmosphere, photons which correspond to the energy levels of the constituent gases are absorbed.

Figure **??** illustrates the processes of creating the three types of spectra, namely the continuous spectrum, absorption line spectrum and emission line spectrum.

As you can see, astronomers can learn a great deal about the nature of an object by looking at its spectrum. Typically, over half the time allocated at an astronomical observatory is devoted to *spectroscopy.* In concept, *spectrographs* are really quite simple: light from a source is sent to a prism (or, more commonly, to a ruled mirror, called a grating) which disperses the light into its component wavelengths (colors). Once the light is dispersed, astronomers then look for any emission or absorption lines that may be present, and begin analyzing the source.

Figure 5: Kirchhoff's laws of spectral analysis.

6 The Lab

This lab will deal with emission spectra. You will be given four different emission spectra taken with the Black Moshannon Observatory spectrograph. One of the spectra will be of a known element (helium), so by finding the location of the emission lines, you will be able to place a "wavelength scale" on the data. You will then turn to the other spectra, which will be of unknown elements. By measuring the wavelengths of the emission lines in these spectra, you should be able to identify the elements.

6.1 Using IDL, the Interactive Data Language

Since this is your first lab with the Sun computers and IDL, take some time to familiarize yourself with the system by trying out some of the commands. Refer to the pages that I gave you last time on using Unix and IDL. Log in and open up another xterm by using the *right* mouse button to bring up a menu and select **xterm** from the **Other tools** submenu. You will start IDL in this xterm, load and run two IDL programs which will read and display your spectra, and then perform a series of wavelength measurements. By doing this, you will be able to create a wavelength solution for your spectra, and find the transformation between pixel value (actually, array element) and wavelength. Once this is done, you will be able to figure out where the prominent emission lines in the unknown spectra lie (*i.e.,* at what wavelength) and determine what element is doing the emitting. Remember, that if anything goes wrong while running an IDL program, type the magic incantation **retall**.

6.2 Procedure

1. Make your own copy of two programs *lab2.pro*, and *pixel.pro*. To do this, type

   cp /afs/psu.edu/users/m/a/mab22/lab2/lab2.pro lab2.pro

   followed by

   cp /afs/psu.edu/users/m/a/mab22/lab2/pixel.pro pixel.pro

   Note: **cp** (copy) is a unix command. If you wanted to execute this, or any other unix while in IDL, you must begin the line with a **$**.

2. Now, enter IDL by typing **idl** (small letters) from your xterm. Remember, most instructions associated with SUN workstations are case sensitive — capital letters are different from lower case letters.

3. Once in IDL, read your four spectra in IDL arrays. This can be done most easily by typing

   .run lab2

BE PATIENT – it may take a few seconds to run! (**lab2** is a very simple program that loads in all the spectra for the lab. It's fairly easy to understand; look at the program by typing

$ page lab2.pro

(The dollar sign is needed since this is a unix, not an IDL command.) Especially note how IDL is told which data files to read, and how the arrays were declared. (You may want to use this construction sometime).

4. To see the arrays which hold your spectra, type **help**. The arrays are named *he, u1, u2,* and *u3.* (In the clever notation of the lab, this stands for helium, unknown #1, unknown #2, and unknown #3.) You can list any of these arrays by typing

   print, he

   (or **u1, u2,** or **u3**, but since there are alot of numbers, they will scroll by very rapidly. Alternatively, you can print only some of the numbers, say, array elements 1001 through 1101, by typing

   print, he(*,1000:1100)

   NOTE THE CONSTRUCTION. The argument **he** is a two-dimensional array, consisting of two columns of numbers of 2560 values each. The first column contains the x-values (pixels); the second column has the list of y-values (counts). The above **print** command says, "print all the columns (note the asterisk wild-car construction), but only rows 1001 through 1101 in each." (IDL uses zero-indexing, which means the first element in the array is element 0, the second is element 1, etc.) This type of command, allows you to operate on only parts of arrays, and is one of the best features of IDL.

5. Make a plot the helium spectrum on the screen. To do this, type

   plot, he(0,*), he(1,*)

   This says, "make a plot using all the elements in the first column of **he** as the x-values and the all the elements of the 2nd column of **he** as the y-values." If you wish to adjust the scales of plot, say, by making the full range of the plot go from −100 to 2000, you can do this by typing

   plot, he(0,*), he(1,*), yrange=[−100,2000], ystyle=1

   If you wish to label the x-axis, you can then add

   plot, he(0,*), he(1,*), yrange=[-100,2000], ystyle=1,xtitle="Pixels"

   A y-axis label can be added similarly. Finally, if yo wish to print out a "hardcopy" of your plot, type

   set_plot, 'PS'

   This will send ALL SUBSEQUENT plotting instructions to a file, called *idl.ps*, INSTEAD of the terminal. You can then make your plot, and when you're finished type

   device, /close

You should then go back into terminal plotting mode by typing

set_plot, 'X'

When this is done, make your hardcopy with the command

$ lpr idl.ps

You should then clean up your directory by deleting the postscript (plot) file that you made with the command

$ rm idl.ps

6. Now introduce IDL to your second program. To do this, type

   .run pixel

   Nothing will happen, but this command lets IDL know that your program exists. This is the program that will do your measurements.

7. You are now ready to begin the lab. Your initial task is simple (but easier said than done). At the back of this lab write-up is a list of the principle emission lines of helium. Look at the lines on your screen, look at the table, and figure out which emission lines correspond to which wavelength. The program **pixel** was written to help you do this (see next step).

8. To measure the position of the emission lines, first start by typing

   pixel, he

   This will plot the entire spectrum and allow you to measure rough pixel values for the emission lines. If you wish to zoom in on a section of the plot and do things more precisely, type

   pixel, he, *x1, x2*

   where *x1* and *x2* are the first and last pixels to be plotted. Record these pixel values (either in a notebook, or with an editor on the computer). Now try to figure out which line corresponds to what wavelength. You can use the fact that the spectral *dispersion,* defined as $(w2 - w1)/(x2 - x1)$, where *w1* and *w2* are the wavelengths which correspond to pixels *x1* and *x2*, changes slowly and regularly across the spectrum.

9. Once you have all the lines correctly identified, you'll be able to find the wavelength of any pixel at all. One way to do this is by simple interpolation between the known lines. Another way (which is the method preferred by astronomers), is to solve the problem mathematically. In essence, what you have a set of wavelengths, $w_1, w_2, \ldots w_n$, which correspond to measured pixels, $x_1, x_2, \ldots x_n$. From these data, you would like to find a function, f, such that $w_i = f(x_i)$. Such a function would allow you to calculate the wavelength for any and all pixels in the data files. For our purposes, a reasonable function to use is a polynomial, $f(x) = a_1 + a_2x + a_3x^2 + \ldots$. The problem, of course, is to find the coefficients. In the language of statisticians, this is known as a problem in regression and methods to solve regression problems are found in many mathematical textbooks. Fortunately, you don't have to research the topic here — IDL provides a function that will perform a polynomial regression. To find the

coefficients of the polynomial, load the pixel values you measured (x) and their corresponding wavelengths (w) into two arrays,

$$\textbf{pixvals} = [x_1, x_2, x_3 \ldots x_n]$$

$$\textbf{wavevals} = [w_1, w_2, w_3 \ldots w_n]$$

then execute IDL's polynomial fitting program by typing

coef = poly_fit (pixvals, wavevals, n)

where n is the degree of the polynomial (2 means quadratic, 3 means cubic, *etc.*). If you then type

print, coef

you will see the coefficients (constant term first). This is your "wavelength solution" which you can use to go from pixels to wavelengths. (Hint: to check the accuracy (and sanity) of your solution, use your polynomial to re-compute the wavelengths of the "known" emission lines. Presumably, the computed values should be close to the actual wavelengths; the difference between the two is an estimate of the error is in your fit. You may also want to use the polynomial to find the wavelength of every pixel in your spectrum and plot pixel number against wavelength.)

10. The other unknown spectra were taken with the same instrument setup, to the wavelength solution will apply to them as well. Use **pixel** and your wavelength solution to find the emission line wavelengths in the unknown spectra and try to identify the elements from the linelists provided.

11. When you're through with IDL, type **exit** to get back to unix. You can logout from there using the mouse.

7 The Helium Comparison Lines

```
THE HELIUM SPECTRUM

  3187.74    He I
  3888.65    He I
  3964.73    He I
  4026.19    He I
  4471.48    He I
  4713.14    He I
  4921.93    He I
  5015.67    He I
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5875.62    He I
6678.15    He I
7065.19    He I
7281.35    He I
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