

Properties of Light and the Atomic Spectra

Objectives: To explore the nature of light and to become acquainted with *spectroscopy* a very useful analytic tool in astronomy, physics and chemistry.

Equipment: Spectrometers, gas tubes, tube power supplies.

Discussion

What is light? From the time of Issac Newton (1642 - 1727) to the beginning of the 20th century the fundamental nature of light was a topic of hot debate among scientists. Newton proposed in his book *Opticks*, published in 1704, that light consisted of a stream of particles. This was known as the *corpuscular* theory of light, and with it Newton was able to explain many common phenomena involving light such as reflection and refraction. Most scientists of this period accepted the corpuscular theory, both because it seemed to explain known phenomena and because Newton was its architect. During Newton's lifetime, however, Dutch physicist and astronomer, Christian Huygens (1629 - 1695), proposed an alternative theory that also explained reflection and refraction. In Huygens's theory light consisted of waves. Wave theory was also useful in explaining a phenomenon known as *diffraction*, or the bending of light around sharp edges. In spite of its successes, Huygens's theory did not receive wide acceptance. All waves known to scientists at the time (sound, water, etc.) traveled through a medium of some sort. Light, on the other hand, travels to us from the sun, moon and stars through the vacuum of space.

In 1801, Thomas Young conducted an experiment that demonstrated another phenomena that could only be explained by the wave theory of light, that of *interference*. Interference occurs when light waves combine in such a manner that they either add together to form a bigger wave or cancel each other out. Figure 1 shows the case of both *constructive* and *destructive* interference. In constructive interference identical light waves that are *in phase* with (i.e., the crests and troughs of each wave correspond with the other) add in such a manner that the combined wave has an *amplitude* (A) that is twice that of either individual light wave. In destructive interference waves that are exactly out of phase (i.e., the crest of one wave corresponds with the trough of the other) cancel each other out when added together. Such behavior could not be explained by the corpuscular theory since at the time no known mechanism could explain how particles could combine so as to cancel each other out. Later work by Maxwell demonstrated that light was actually a high frequency electromagnetic wave (like radio or television waves) that traveled through space with a speed of 3×10^8 m/s. By the dawn of the 20th century, the wave theory of light was firmly entrenched in scientific dogma.

As resounding as the evidence for the wave behavior of light was, there were still experiments that gave results that could not be explained by the wave theory. One such result is the *photoelectric effect*. The photoelectric effect is a phenomenon whereby electrons are ejected from a metallic surface that has been exposed to light. This phenomenon cannot be reasonably explained with the wave theory.

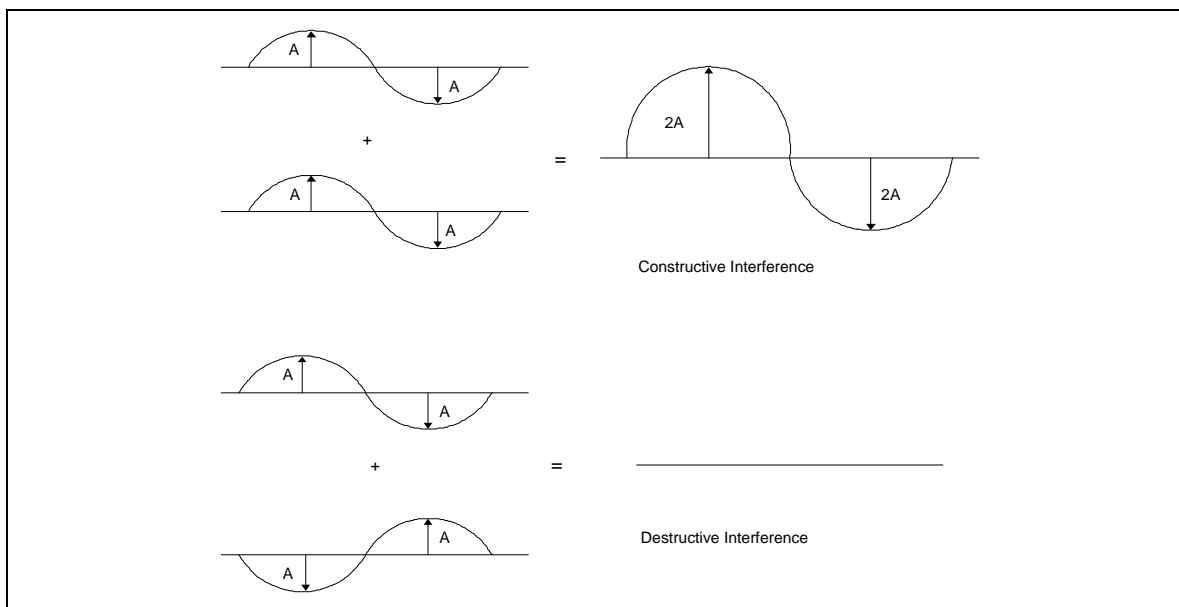


Figure 1. Interference in waves.

An explanation of the photoelectric effect was proposed by Albert Einstein in 1905. Einstein, drawing upon the earlier work of Max Planck (1858 - 1947) proposed that light was composed of discrete bundles of energy called *photons* and that the energy (E) of each photon was proportional to the frequency (f) of the light wave:

$$E = hf$$

where $h = 6.63 \times 10^{-34} \text{ J} \cdot \text{s}$ is *Planck's constant*. Notice that Einstein's explanation of the photoelectric effect contains elements of both the particle (energy of a single photon) and wave (frequency) theories of light.

This is the essential nature of light. At times it exhibits particle-like properties and at other times it exhibits wave-like properties. The best answer to the puzzle "What is light?" is that it is both particles and waves: sometimes the particle properties dominate and sometimes the wave properties dominate.

For the purposes of this procedure we may assume that the wave properties of light dominate. The wavelength of light (λ) is related to its color. In the visible spectrum which extends from approximately 400 - 700 *nanometers* (a nanometer is 10^{-9} meters), longer wavelengths correspond to red and orange and shorter wavelengths correspond to violet and blue. You have probably heard about *ultraviolet* light and its associated dangers in relationship to the depletion of the earth's ozone layer. This is light with wavelengths in the 100 to 400 nanometer range. Wavelengths from 1 millimeter to 10 microns (10^{-3} to 10^{-7} meters) are known as *infrared* light.

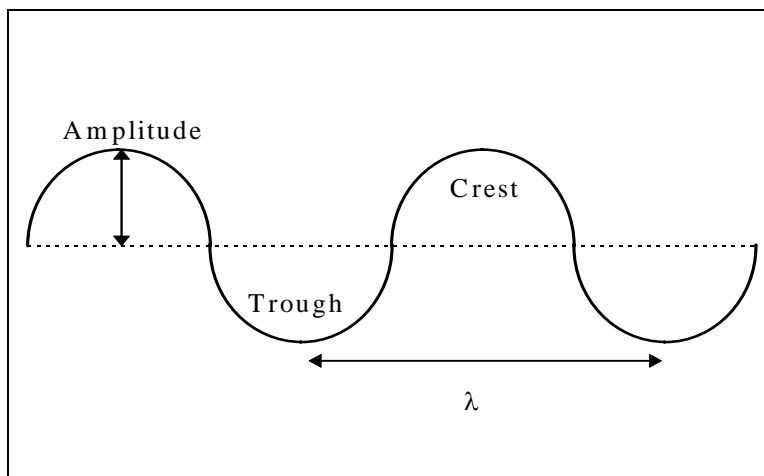


Figure 2. Characteristics of a light wave.

The frequency of a light wave is equal to the speed of light divided by its wavelength:

$$f = \frac{c}{\lambda}$$

Since the speed of light in free space is constant ($c = 3 \times 10^8$ m/s) the longer the wavelength the lower the frequency. All visible light waves have frequencies of around 10^{15} Hz.

Spectroscopy

Spectroscopy is an extremely useful analytic technique that involves analyzing the light given off by an object. A spectrometer is a device that uses diffraction and interference to separate the light given off from an object into different colored lines known as a spectrum. Each spectral line corresponds to light of a different wavelength. Every element in the universe (e.g. Hydrogen, Helium, etc.) has a unique spectrum. This is possible because of the unique electronic structure of each element. When electrons in atomic orbitals gain energy either through collisions with other electrons or by exposure to electromagnetic waves (such as light), they are promoted to a higher orbit corresponding to greater energy. Normally an electron immediately decays back to its original lower-energy orbital emitting energy that is *exactly* equal in difference in the energy between the two levels. For the most part this energy is emitted by photons some of which are in the visible spectral region. Since energy levels are different in each element the corresponding spectral patterns are also different in each element. By carefully analyzing the spectra given off by an object such as a star, one may gain insight as to its composition.

Let us consider the various types of spectra and their origin. There are three types of spectra: (1) bright line or emission line spectra, (2) continuous spectra, (3) dark line or absorption line spectra.

A bright line or emission spectra is produced by a glowing gas which radiates energy at specific wavelengths characteristic of the element or elements that make up the gas. The spectrum consists of a number of bright lines against a dark background. Each bright line has a color that represents a specific wavelength.

A continuous spectrum is produced by a glowing solid, liquid or gas under certain conditions. A prism, for example, forms a continuous spectrum when white light is transmitted through it. The spectrum appears as a smooth transition of all colors in the visible spectrum from the shortest or the longest wavelength without any gaps between the colors.

A dark line or absorption spectrum is produced when a cooler gas lies between the observer and the object emitting a continuous spectrum. The cooler gas absorbs specific wavelengths of light passing through it. The wavelengths absorbed are determined by the elements that compose the gas. Since no two elements absorb the exact same wavelengths, it is possible to determine the elemental composition of the gas by examining the spectra. A dark line or absorption spectrum appears as a continuous spectrum of all colors with a number of dark lines through it. If the dark lines are closely spaced in some parts the clumps of dark lines are known as bands.

Procedure

You are to observe the emission spectrum (bright line spectrum) of Hydrogen, Helium, Sodium and Mercury. You will do this with the aid of a spectrometer such as the one shown in Figure 3. The emission spectra of each element will be produced by placing a tube containing the element into a device that sends an electrical current through the tube. This electrical current will cause the material in the tube to glow, thus producing the emission spectra.

Create a table in your lab notebook with the headings: Line #, Color, θ_L , θ_R , θ_{AVE} , λ . Your lab instructor will show you how to put a gas tube in the high voltage source and turn it on. When switching tubes, wait for the tube to cool after turning it off before touching it. Do not attempt to change tubes with the power supply turned on! Begin with the Mercury (Hg) tube.

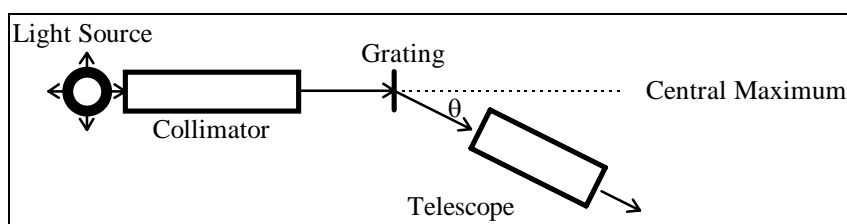


Figure 3. Schematic of a spectrometer.

Dim the room lights and look into the telescope of the spectrometer. Align the telescope and the collimator on the spectrometer so that they are in a straight line. Adjust the position of the high voltage source until the narrow portion of the tube is directly in front of and centered on the narrow slit at the end of the collimator. Peer through the telescope and adjust the width of the slit to produce

a bright, narrow line the same color as the glowing gas in the tube. You will have to experiment with moving the light source around a bit to get the brightest line in the telescope. This central bright line is known as the central maximum. It is not a spectral line but its position is important. Note its position on the vernier scale at the base of the spectrometer in your lab notebook. The central maximum is (or should be) located at about 180 degrees on the vernier scale.

Move the telescope to the left until you encounter a violet line. This is the first spectral line in Mercury. Center the telescope as best you can on this line and note its position on the vernier scale. If the line is at, for instance, 175 degrees, it is said to be deviated 5 degrees from the central maximum ($180^{\circ} - 175^{\circ} = 5^{\circ}$). Record this angle of diffraction under θ_L . Continue scanning left noting the position and color of each line that you encounter. You should encounter blue, green yellow, orange and possibly a faint red line. After passing the red line at the far end of the spectrum you will notice that if you keep scanning left you eventually encounter another violet line just like the first but fainter. This is the beginning of the second order ($n = 2$) spectrum. Stop scanning at this point and return the telescope to the position of the central maximum.

Scan to the right this time and note the position of each line that you encounter. Notice that you encounter the same sequence of lines at about the same relative positions from the central maximum. If you locate a blue line at 185 degrees the line is said to be 5 degrees from the central maximum ($185^{\circ} - 180^{\circ} = 5^{\circ}$). Record these angles under the heading θ_R . Next record the average diffraction angle, θ_{AVE} for each line. This may be computed as follows:

$$\frac{\theta_L + \theta_R}{2} = \theta_{ave}$$

The wavelength for each line in the first order ($n = 1$) may be then computed by using the following formula:

$$n\lambda = d \sin \theta_{ave}$$

Where d is the width of the lines on the diffraction grating in the spectrometer that you are using. In order to compute d , first examine the diffraction grating. The grating is a slide with a number of grooves etched into its surface. The number of lines per cm (or lines per inch) is printed on the slide. In order to calculate the width of an individual line all that is necessary is to take the inverse of this number. Remember to convert to metric units if necessary.

Compute the value of λ for each line in the Mercury spectrum and compare your values with those provided by your lab instructor. Repeat this experiment for Hydrogen and Helium. You will not encounter the same lines in each element! What you will encounter is a unique spectral "fingerprint" for each element that you study.

Recall that the energy of a photon of light is given by $E = hc/\lambda$ where c is the speed of light and h is Planck's constant, 4.14×10^{-15} eV·s. Record the energy of the photons for each line you

have identified in the three spectra you have examined (use the values provided by your lab instructor for this). This data may be used to construct energy level diagrams which we will do for hydrogen. The transitions in the visible spectral region for hydrogen are known as the *Balmer Series*. The lowest non-radiating state in hydrogen (known as the *ground state*) is designated by $n = 1$ and has an energy of -13.6 eV. The next lowest state is -3.40 eV. To calculate the energy levels in hydrogen:

$$E_n = -\frac{13.6}{n^2} eV \quad n = 1, 2, 3, \dots$$

Compute the first 10 energy levels in hydrogen and construct an energy level diagram like the one shown in Figure 5. All of the transitions in the visible spectrum of hydrogen correspond to decays of electrons to the same state from various excited states. Fill in the transitions on your energy level diagram.

For example: The first order red line in the hydrogen spectrum is at $\lambda = 656.3$ nm. This corresponds to an energy of 1.89 eV which is exactly the difference in energy between the $n = 3$ and $n = 2$ states. This line, therefore, corresponds to the transition between the $n = 3$ and $n = 2$ states. Indicate this by drawing a line with an arrow starting at the 3rd excited state to the 2nd.

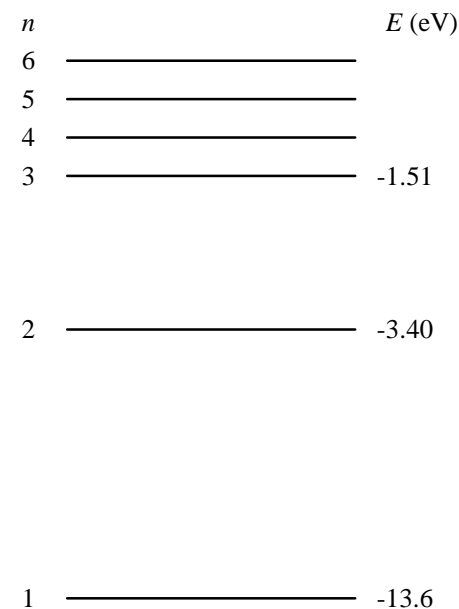


Figure 3.

Exercises

1. What is the frequency of a light wave that has a wavelength is 588 nanometers? What is its color?
2. Compare and contrast constructive and destructive interference.
3. Why are the spectral lines in helium at different wavelengths than those in hydrogen?
4. What is the function of the collimator on the spectroscope?
5. Are any of the transitions to the ground state visible in hydrogen?