

## Light and Spectra

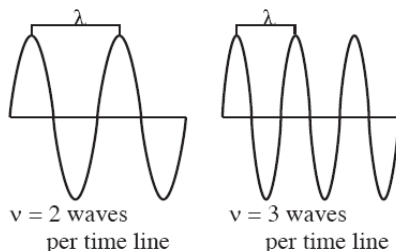
### INTRODUCTION

Light and color have intrigued humans since antiquity. In this experiment, you will consider several aspects of light including:

- The visible spectrum of colors (red to violet)
- Bright line spectra as emitted by an excited gas or solid
- The relationship between color, wavelength, frequency and energy.

### Background and Theory

Light is a form of energy called electromagnetic radiation. It has wavelength and frequency. Wavelength,  $\lambda$  (lambda), is the distance between adjacent wave crests. Visible light has wavelengths in the 400 nm to 700 nm range (1 nm =  $10^{-9}$  m). Frequency,  $\nu$  (nu), tells how many waves pass by a point in a second. Violet light, with a wavelength of 400 nm, has a frequency of  $7.5 \times 10^{14}$  per second, many more waves than are shown in the diagram.



The energy of electromagnetic radiation varies directly with the frequency, and inversely with the wavelength. Thus, violet light is of higher energy than red light, which has wavelengths in the 650 nm range. White light, such as that from an ordinary incandescent light bulb, is a mixture of wavelengths in the visible range. When white light strikes a prism or diffraction grating, the light is dispersed into a **continuous spectrum** of visible colors. A **non-continuous spectrum** occurs when an electric current passes through a gaseous element in a gas discharge tube or when metal ions are put into a flame. This type of spectrum, called a bright **line spectrum**, is not continuous but instead contains only certain colors at particular wavelengths.

A bright line spectrum is due to transitions of electrons between energy levels in atoms or ions. When an atom or ion absorbs energy, such as from a flame or electrical source, it absorbs only certain discrete amounts of energy. These amounts are the difference in energy between lower and higher electron energy levels. When the electrons return to lower energy levels, energy is emitted in the form of a photon. Since atoms have a number of energy levels available, there are a number of different energies that can be absorbed and released. For example, the visible portion of the spectrum for mercury contains the following colors:

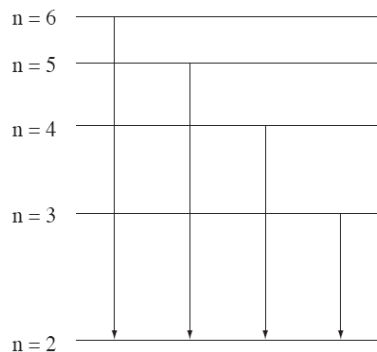
COLOR	$\lambda$ , nm	COLOR	$\lambda$ , nm
violet	405	yellow	579
blue	436	orange	623
green	492	red	689

The unique spectrum given by each element provides a simple means of analysis for elements. After exciting a substance with a spark or in a flame, the position of the spectral lines produced gives the identity of the elements in the substance, and the intensity of the lines gives the quantity of the elements. Of course, instruments that do this have built-in databases that allow

for the experimental data to be compared with previously compiled data.

Also, notice that the lines have different widths. This is typical of spectral lines, and spectroscopists rank lines according to their appearance and abundance. Hence, the designations *s*, *p*, *d*, and *f* are for electron orbitals (from sharp, principal, diffused and fundamental).

The energies of the emitted wavelengths can be correlated with the energies of electron transitions from higher to lower energy levels. During an energy<sub>high</sub>→energy<sub>low</sub> electronic transition, the greater the difference between the energy levels, the larger is the energy released, and the shorter is the wavelength of the photon emitted. In the energy level diagram of a generic atom shown below, the left-hand arrow represents an electron releasing energy as it drops from the 6<sup>th</sup> to the 2<sup>nd</sup> energy level.



The energy is released in the form of a photon. This photon will have more energy (and have a shorter wavelength) than the photon given off in the transition represented by the right-hand arrow. In terms of energy,  $(n_6 \rightarrow n_2) > (n_3 \rightarrow n_2)$ . The larger energy would be in the violet-blue portion of the spectrum. The smaller energy would be in the red-orange portion of the spectrum.

### **Frequency and Energy**

The relationship between wavelength and frequency is:  $\lambda \cdot \nu = C$ , where C is the speed of light,  $3.00 \times 10^8$  m/s. The violet line with  $\lambda = 405$  nm from the mercury spectrum would have a frequency calculated as follows

$$\nu = \frac{C}{\lambda} = \frac{3.00 \times 10^8 \text{ m}\cdot\text{s}^{-1}}{405 \times 10^{-9} \text{ m}} = 7.41 \times 10^{14} \text{ s}^{-1}$$

To find the energy of a photon, the Planck relationship is used:  $E = h \nu$ , where h is the Planck constant,  $6.63 \times 10^{-34}$  J.s. The energy of the violet line is

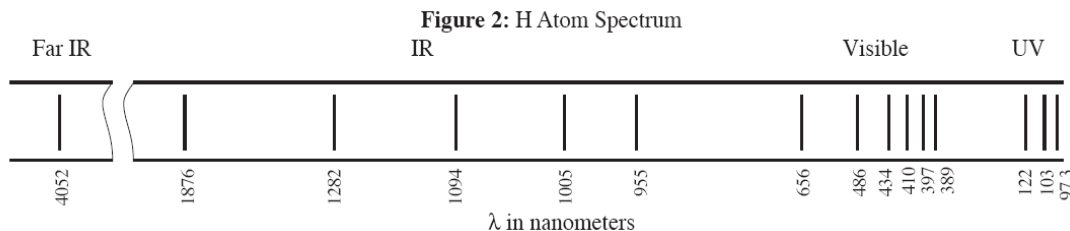
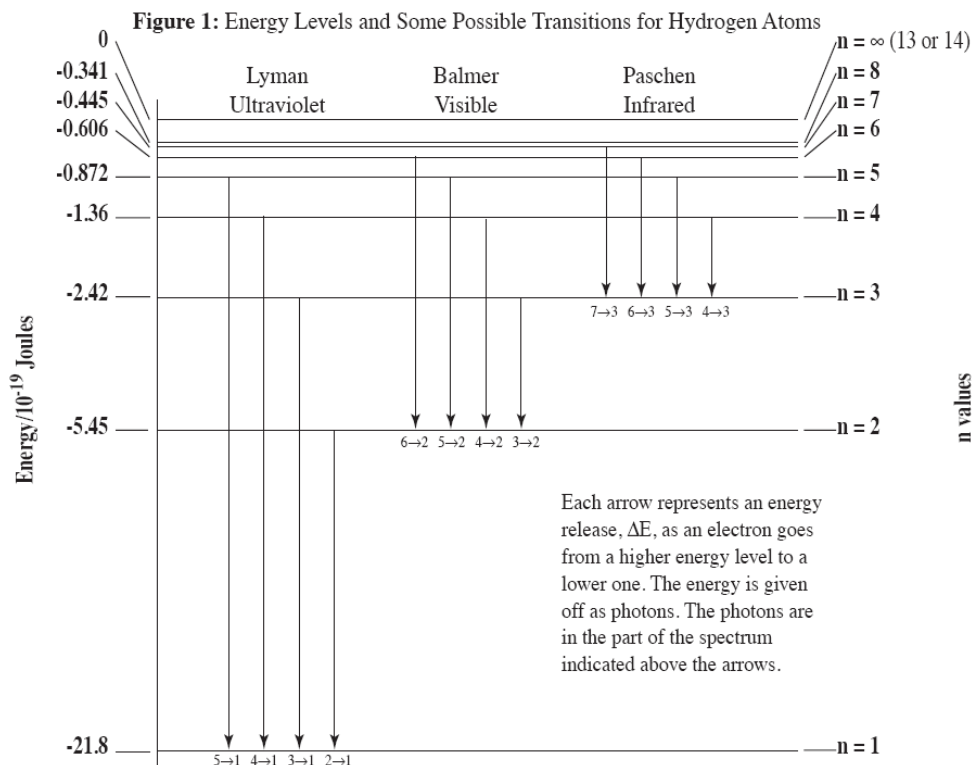
$$E = (6.63 \times 10^{-34} \text{ J}\cdot\text{s})(7.41 \times 10^{14} \text{ s}^{-1}) = 4.91 \times 10^{-19} \text{ J}$$

This is the energy emitted when an electron in one particular energy level in a mercury atom drops down to a lower particular energy level. The energy  $4.91 \times 10^{-19}$  J might seem extraordinarily small, but in terms of a mole of atoms, there would be  $296 \text{ kJ}\cdot\text{mol}^{-1}$ . (Can you verify this?).

This is a good time to make a general observation about quantities. You will hear people say, and perhaps you say, “that seems like an awfully small amount”. The response to this is, “compared to what?”. A mercury atom, one of the really heavy atoms, weighs about  $3 \times 10^{-22}$ g. Compared to your weight, a mercury atom weighs “nothing”. Compared to a neon atom, the mercury atom weighs 10 times as much. A thousand percent heavier! The moral of the story is: Consider everything in perspective.

## Hydrogen Spectrum

Hydrogen atom is the simplest atom and a detailed study of its spectrum leads to valuable insights. Shown in Figure 1 below are the energy levels in the hydrogen atom as calculated from the Schrödinger equation. The numbers on the left give the energy of each level in joules. The number on the right gives the principal quantum number of each level. The names along the top are the names of the scientists who first characterized the spectrum corresponding to these transitions. Shown in Figure 2 is a black and white representation of the experimentally observed spectrum of the hydrogen atom. IR is infrared, VIS is visible and UV is ultraviolet. The wavelength of each line in nanometers is listed below the line. Figure 2, which is experimental, can be calculated from Figure 1, which is theoretical.



According to spectral theory, each line in the spectrum (Figure 2) comes as an electron drops from a higher to a lower energy level. In Figure 1, the various vertical arrows show a number of such possible transitions. The energy of a transition, for example, the one indicated by the left hand vertical arrow in Figure 1 above, can be calculated by subtracting the energy of the initial state from the energy of the final state (found on the left of the diagram):

$$\Delta E_{\text{upper} \rightarrow \text{lower}} = E_{\text{lower}} - E_{\text{upper}} = E_{\text{photon}} = h\nu$$

Hence, the energy of transition  $n_5 \rightarrow n_1$  is

$$\Delta E_{5 \rightarrow 1} = (-21.8 \times 10^{-19} \text{ J}) - (-0.872 \times 10^{-19} \text{ J}) = -20.9 \times 10^{-19} \text{ J}$$

The negative value indicates energy released. Using the relation shown above you can prove that this transition corresponds to frequency  $\nu = 3.08 \times 10^{15} \text{ s}^{-1}$  and wavelength  $\lambda = 97.3 \text{ nm}$ , the far right line in hydrogen spectrum (Figure 2).

## EXPERIMENTAL PROCEDURE

STUDENTS SHOULD WORK INDIVIDUALLY. Do not copy data from another classmate. Obtain your own observations!! Tell your instructor if you have serious eyesight problems or if you suffer from color blindness. Data collection for this experiment is simple. However, understanding what is going on can be quite difficult. Take your time to observe things clearly and make detailed drawings so that you can extract numerical information from them to complete your lab report.

THE LIGHTS IN THE ROOM MUST BE TURNED OFF for the duration of this experiment. Some windows might also need to be covered up.

### *Using a spectroscope*

Spectroscope contains a diffraction grating, which functions similarly to a prism, as it can take light and separate it into its different wavelengths (like a water does to produce a rainbow). The process is known as diffraction. Diffraction is the bending of light by an edge or a slit. The grating in spectroscope has hundreds of grooves that serve as slits. The angle at which the light bends depends on its wavelength. A prism actually causes the light to be refracted. The light travels through the prism and comes out the other face of the prism. Different wavelengths "bend" differently inside the prism so that the light that comes out of the prism is separated into its component wavelengths. Because we are using our eyes as the detectors, we can only observe the visible components of the light. The light source we are using will emit in other regions of the electromagnetic spectrum such as the infrared and the UV.

Handle the spectroscope with care. It can break if dropped. Look in through the narrow end of the spectroscope. You should see a little slit (opening) towards the left, and the scale on the right side. You might also see colors depending on how much light is coming into the spectroscope. If you have very poor eyesight, you might not be able to see the scale clearly. The scale is in nanometer unit and at 10 nm increment. To analyze the light output from a lamp source you will need to aim the spectroscope at the source and be able to see light from the source entering through the slit. For a light bulb or a fluorescent bulb, you can see the spectrum in the spectroscope even if you are four feet away from the source. For a less intense source you might need to get as close as one foot or so from the bulb.

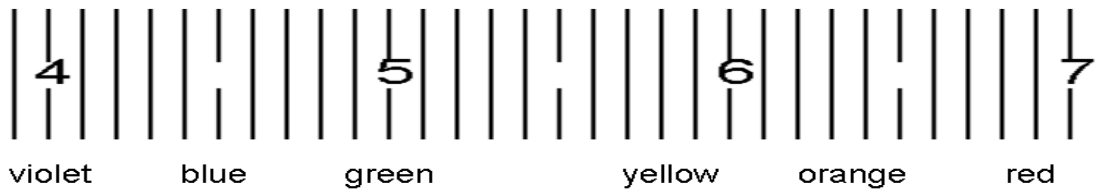
### *Continuous Spectrum*

Direct the spectroscope to the sunlight. The theory available 100 years ago predicted that as the temperature increased, the spectrum would completely shift over into the ultraviolet, with no reds or yellows or greens remaining. This inadequate theory had physicists mumbling about the "ultraviolet catastrophe". Theoretical physicists looked to develop theories consistent with experimental data. Max Planck developed a new theory which has held up for close to a century. The equation:  $E = h\nu$  is part of the theory.

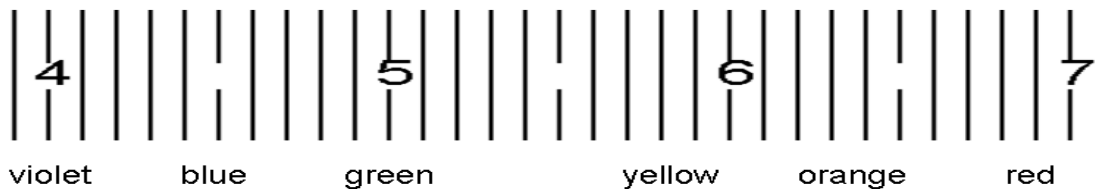
### *Line Spectra*

The conditions used in this part of the experiment, namely, high voltage through a low pressure sample of an element, produce a characteristic line spectrum for each element. Each element is labeled and set-up properly. Plug in the high voltage source if necessary.

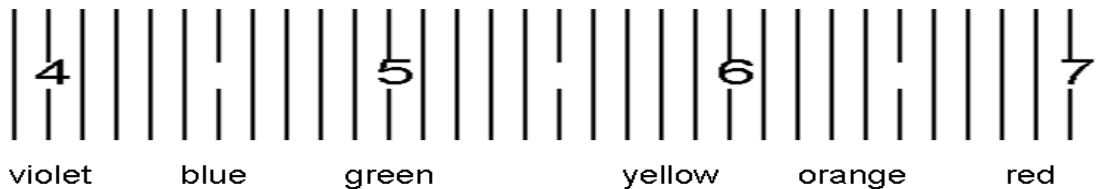
**Hydrogen Spectrum:** Using your spectroscope look at the light being emitted by the hydrogen lamp. The closer you get, the clearer it appears. Observe the lines and draw them on the chart below.



**Mercury Spectrum:** Using your spectroscope observe the lines and draw them exactly on the chart below. Just as H and Hg differ, so each atom has a unique spectrum.



**Neon Spectrum:** Using your spectroscope observe the lines and draw them exactly on the chart below. Notice the number of lines in the red and yellow portion of the spectrum.



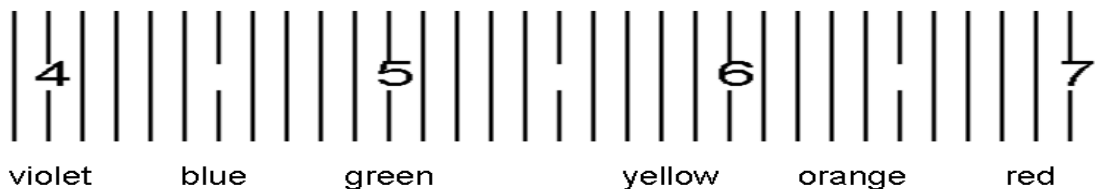
## Light and Spectra

Name \_\_\_\_\_

Date \_\_\_\_\_

Partner's Name \_\_\_\_\_

**Hg Spectrum.** The one you observed in the lab is a low resolution spectrum. The spectral charts posted in your textbook show higher resolution spectra of a number of substances. Look at a chart and sketch the spectral lines of mercury roughly to scale in the strip below: (numbers are  $\lambda \times 10^2$  nm)



Notice that the spectrum has three lines shown in the yellow region. Record the actual wavelengths from the chart, then calculate the frequency and energy of each of these three lines in the mercury spectrum:

Wavelength (nm)	Frequency ( $s^{-1}$ )	Energy (J)
_____	_____	_____
_____	_____	_____
_____	_____	_____

**Hydrogen Spectrum.** For hydrogen atom, the energy of a photon associated with a particular energy level ( $n = 1, 2, 3, \dots$ ) is  $E_n = \frac{-2.18 \times 10^{-18} J}{n^2}$ .

1) Calculate the energy for the first five electron energy levels of hydrogen atom. Those for  $n = 1$  and  $n = 5$  are done.

$E_n$ (in J)	$E_n$ (in J)
$n = 1$ $-2.18 \times 10^{-18} J$	$n = 4$ _____
$n = 2$ _____	$n = 5$ $-8.72 \times 10^{-20} J$
$n = 3$ _____	$n = 6$ _____

2) Fill in the energy values for the possible transitions. This will give you the energy of the photon emitted when that particular transition occurs. Then calculate the energy gap  $\Delta E$  corresponding to each of these transitions.

$\Delta E$ (in J)	$\Delta E$ (in J)
$n = 6 \rightarrow n = 1$ _____	$n = 5 \rightarrow n = 1$ $-2.09 \times 10^{-18}$ J
$n = 6 \rightarrow n = 2$ _____	$n = 5 \rightarrow n = 2$ _____
$n = 6 \rightarrow n = 3$ _____	$n = 5 \rightarrow n = 3$ _____
$n = 6 \rightarrow n = 4$ _____	$n = 5 \rightarrow n = 4$ _____
$n = 6 \rightarrow n = 5$ _____	
	$n = 3 \rightarrow n = 1$ _____
$n = 4 \rightarrow n = 1$ _____	$n = 3 \rightarrow n = 2$ _____
$n = 4 \rightarrow n = 2$ _____	
$n = 4 \rightarrow n = 3$ _____	
	$n = 2 \rightarrow n = 1$ _____

3) Calculate the energy corresponding to each of the color lines in hydrogen spectrum that you observed and recorded. Determine which transitions are closest in energy values to those calculated in section 2) above.

Wavelength (nm)	Frequency ( $s^{-1}$ )	Energy (J)	Transition
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

### Post-Lab Questions and Exercises

**(All questions must be answered during the lab and submitted with your lab report at the end of the lab period).**

*Please answer the following questions and show all work and units. Express all answers to the correct number of significant digits.*

- 1) When looking at the hydrogen light directly with your eyes, does it look white? If not, what color does it exhibit? Can you explain it?
  
  
  
  
  
  
  
  
  
  
- 2) When looking at the mercury light directly with your eyes, does it look white? If not, what color does it exhibit? Can you explain it?
  
  
  
  
  
  
  
  
  
  
- 3) When looking at the neon light directly with your eyes, does it look white? If not, what color does it exhibit? Can you explain it?
  
  
  
  
  
  
  
  
  
  
- 4) Under what conditions did you see line spectra (very specific wavelengths)? Under what conditions did you see band spectra (a range of wavelengths)?
  
  
  
  
  
  
  
  
  
  
- 5) For hydrogen spectrum, to which series of transitions do the lines you observed belong?