

Experiment 13: Emission and Absorption Spectroscopy

Purpose

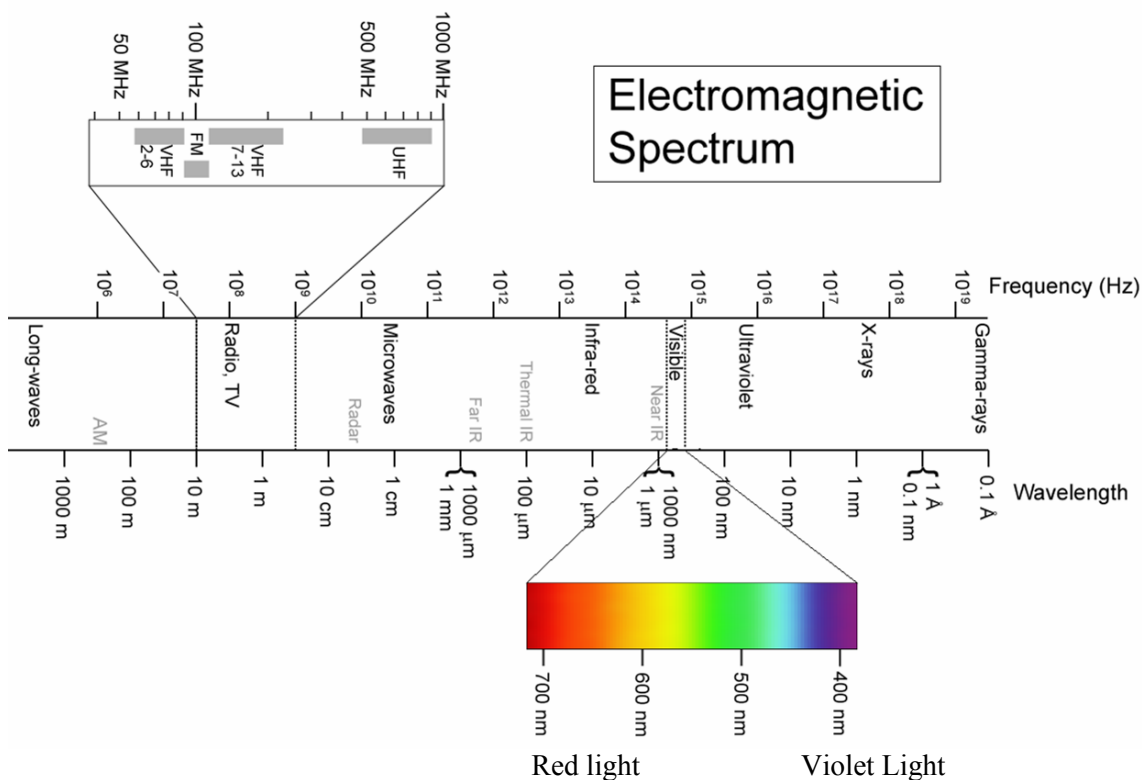
In this experiment you will build and calibrate a spectroscope, measure the wavelengths of spectral emission lines viewed with the spectroscope, and observe the relationship between emission and absorption.

Discussion

Light (electromagnetic radiation) is transmitted through space as a wave. All light travels at the same speed ($c = 3.00 \times 10^8$ m/s) but varies in wavelength and frequency. For a given energy of light, the product of wavelength (λ) and frequency (ν) is equal to the speed of light:

Equation 1
$$c = \lambda \nu$$

Light can be broken down into regions called the electromagnetic spectrum. Depending on the wavelength and frequency of the light, it will fall into one of the following regions. Notice that the visible part of the spectrum can be broken down into colors which corresponds to different wavelengths and frequencies as well.



In 1900, Max Planck proposed the idea that light is "quantized," that is, it consists of tiny packets called photons. The energy of a single photon of light is directly proportional to its frequency.

Equation 2
$$E = h\nu$$

where E is energy in joules, ν is frequency in cycles per second, and h is Planck's constant which has a value of 6.63×10^{-34} J·s.

These equations can be combined to make a third useful equation which relates the energy of a photon of light to wavelength:

Equation 3

$$E = \frac{hc}{\lambda}$$

According to the Bohr Model of the atom, if electrons are "excited" to a higher level by flame or voltage, these electrons can then emit photons of light in order to return to the lower energy orbital. The energy, and thus the wavelength, of the photons emitted will depend on the energy gap between levels. The result is an emission "line". An emission spectrum records all such lines for a substance. For example, when gaseous elements are heated, they emit light as characteristic spectral lines with certain intensities and wavelengths. Emission lines for metal ions in salts are visible when the salt (either solid or in solution) is placed in a flame. The resultant unique emission spectra have been used to identify elements for more than a century. Before there was any explanation for the emission spectra, physicists had found empirically (purely from experimental data) that the wavelengths of the emission lines for hydrogen fit the equation:

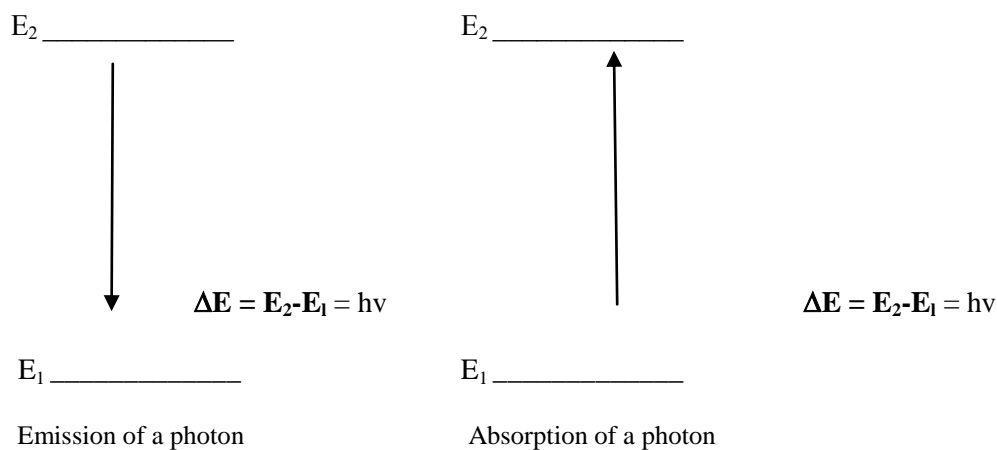
$$\text{Equation 4} \quad \frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

where λ = wavelength of a given spectral line
 $R_H = 1.097 \times 10^7 \text{ m}^{-1}$ (the Rydberg constant)
 $n_1 = 1, 2, \text{ or } 3$
 $n_2 = \text{an integer larger than } n_1$

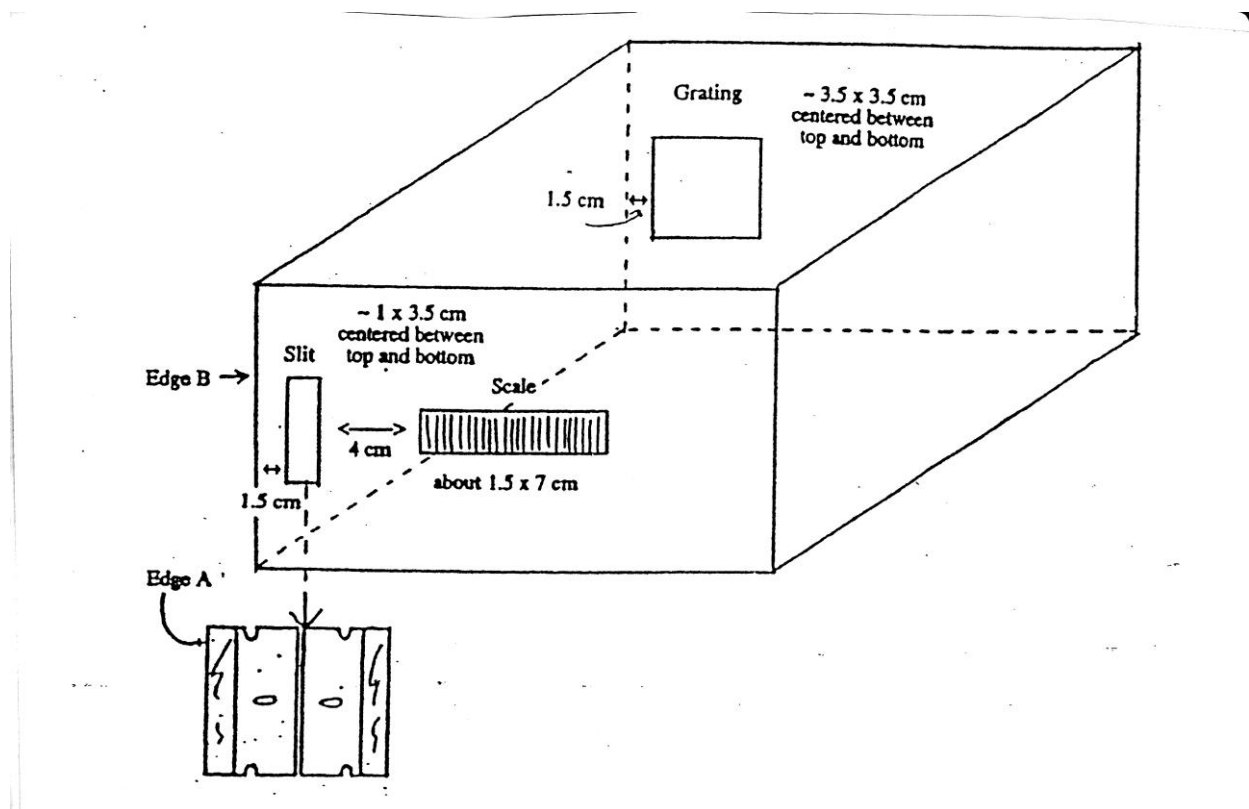
Niels Bohr was the first to propose that photon emission was caused by the transition of an electron from a higher energy state to a lower one, and that n_2 and n_1 in the equation above correspond, respectively, to the quantum numbers of the higher and lower energy states.

Similarly, in order to move to a higher energy level, an electron must absorb an amount of energy that matches the energy "gap" between the two levels. This energy can be supplied by flame, high voltage from an electric arc, or by photons of the proper frequency.

An absorption spectrum records the wavelengths of light absorbed by a substance when its electrons move from lower to higher energy levels. When viewing an object that has absorbed some light, your eye receives only the light that was reflected rather than absorbed. Thus, an object that absorbs light in the yellow region of the spectrum leaves mostly green, blue, and violet for your eye. The object looks blue, which is yellow's complementary color.



In this experiment you will build your own spectroscope, a device to observe spectra, from a cardboard box. With it, you will measure and record both absorption and emission data. A diagram of the scope is shown below.



To use the scope, you look into the grating (a device that works like a prism) and align the slit with a light source. Light enters the slit and strikes the grating, which is nearly transparent. Most of the light passes through the grating, but part of it undergoes diffraction: it is broken up into its component wavelengths, like a rainbow, and projected back onto the scale. Through the grating you will see spectral lines superimposed on the scale, from which you can read their relative wavelengths.

Precautions

Discharge lamps operate at **high** voltage with exposed electrodes. Do not touch them. The mercury and hydrogen lamps emit some ultraviolet light. Protect your eyes by viewing through the glass door of the fume hood.

Procedure

A. Spectroscope Construction (Optional)

Construct a spectroscope using the directions below or those found at http://spaceplace.nasa.gov/en/educators/tes_spectroscope.pdf which use a DVD as the prism.

1. Obtain a box, a grating, two razor blades, and a piece of graph paper for the scale.
2. Refer to the drawing above for proper spacing of holes. Carefully cut holes in the box for the grating, blades and scale. Fold the box into shape and tape it closed.
3. Tape the razor blades over the slit hole, with edge A of the blade as close as possible to edge B of the box, and the sharp edges of the blades not more than 1 mm apart. Don't use too much tape yet as some later adjustment may be needed.

4. Darken the major vertical lines on the graph paper with a pen or pencil, write the numbers from zero to 6 on alternate vertical lines, and tape the scale in place over the appropriate hole so that it faces you when viewing through the grating hole. Make sure all the markings are visible!

5. Tape the grating loosely over its hole (handle the grating by the edges) and look through the grating and slit at a fluorescent light. If a rainbow spectrum is superimposed on the scale, tape the grating firmly in place. (Note: You may need to rotate the grating a few degrees until the spectrum is completely superimposed on the scale). If the rainbow appears directly above and below the slit, rotate the grating ninety degrees before taping it firmly.

B. Spectroscope Calibration

6. Look through the grating and slit at a mercury vapor lamp. You should see three bright lines: yellow, bright green, and blue. More faint lines will probably be visible as well. Read and record the scale positions of the brightest mercury emission lines as precisely as possible. Since the lines will shift slightly as you move the box, try to be consistent in where you hold the box.

7. Make your calibration graph in Excel by enter the wavelengths in column A (x-axis) and corresponding scale readings in column B (y-axis). Complete the calibration graph by include labeled axes with units and a graph title. Add a trendline to your graph. We will use the relationship of the trendline to determine the wavelength (x) of light produced by other samples based on where the emission lines appear on the spectroscope scale (y). If your data points do not form a straight line, take new readings to obtain a good calibration graph.

C. Spectral Analysis

8. Gas discharge: Read and record the scale positions and colors of the emission lines from the hydrogen lamp. You should see at least three lines. Repeat for the helium lamp, which will have several lines.

9. Flame tests: For this test, you will place small quantities of crystals from three different salts on the tip of a nichrome wire probe. In order to get them to stick to the wire, you may need to wet the wire. Beginning with the NaCl sample, position the burner so that the flame's blue cone engulfs the salt. Look at the resultant colored flame through your spectroscope by positioning the slit close to the flame. Record the scale position and color for each observed emission line. Repeat this procedure for LiCl and SrCl₂. (Note: It is advisable to begin with NaCl since sodium is a common contaminant.)

10. Absorption: Look at the spectrum from an ordinary light bulb through your scope. Then place a small beaker containing potassium permanganate solution between the light bulb and your scope. Record the wavelength region most strongly absorbed by the sample.

Emission and Absorption Spectroscopy Data

Name _____
Date _____

Gas Discharge Tubes

	Color	Scale Reading	Wavelength (nm)
Mercury	blue		435.8
	green		546.1
	yellow		579.0

	Color	Scale Reading	Wavelength (nm) determined from calibration graph
Hydrogen			

	Color	Scale Reading	Wavelength (nm) determined from calibration graph
Helium			

4. Flame emission (Note: Sodium, a common contaminant, often produces its emission line in the spectra of the other salts).

	Color	Scale Reading	Wavelength (nm) determined from calibration graph
NaCl			
LiCl			
SrCl ₂			

5. Absorption

Color of KMnO₄ solution _____

KMnO₄ wavelength region absorbed:

Scale Reading: _____ to _____

Wavelength (from graph) _____ nm to _____ nm

Results and Questions

1. Use your calibration curve to determine wavelengths (in nm) for each emission line recorded for the hydrogen, helium, and the salts. Record these values on the data page where indicated, and use the correct number of significant figures. Use the scale to determine the wavelength

region absorbed by KMnO_4 and record on the data page where indicated. Be sure to attach the calibration graph to your report

2. For hydrogen, calculate the emission wavelengths (in nm) for the electron transitions described below, using the Bohr theory equation

$$\text{Equation 4} \quad \frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

Make the closest match for your experimental values for hydrogen with these calculated values. **Show a sample calculation below.** (Note: Although you need to calculate values for all four transitions, you only observed three of them so one of the lines in the observed wavelength column will be left blank)

Sample Calculation:

n_2	n_1	Calculated wavelength	Experimentally observed wavelength
6	2	_____	_____
5	2	_____	_____
4	2	_____	_____
3	2	_____	_____

3. For helium, match your measured wavelengths with the published handbook values below. You will have observed the most intense lines, in the lab, so they should be given priority when matching yours with those listed.

Wavelength (nm)	Intensity	Observed Wavelength (nm)	Wavelength (nm)	Intensity	Observed Wavelength (nm)
388.9	1000		471.3	40	
396.5	80		492.1	50	
400.9	10		501.6	100	
402.6	70		504.8	15	
412.0	35		541.1	50	
414.4	15		587.56	1000	
438.8	30		587.59	10	
443.8	10		656.0	100	
447.1	100		667.8	100	
468.6	300		706.5	70	

4. For sodium ion, calculate the percent error between your measured wavelength and the handbook value of **588.9** nm.

$$\% \text{ error} = \frac{\text{observed} - \text{true}}{\text{true}} \times 100$$

5. KMnO_4 solutions are magenta/purple in color, with the depth of the color depending on the concentration (very concentrated solutions may look almost black.) Based on your data for the wavelengths absorbed by the solution, explain why it is purple. Careful! Remember that the process responsible for the color of the solution is different from the process that is responsible for the colors of the gas discharge tubes and flames tests.

6. The amount of energy required to completely remove an electron from a hydrogen atom can be found by assuming it is being taken from the $n=1$ level to the $n= \infty$. Calculate this amount of energy.

