Spectroscopy

Goals:
Observe and describe emission spectra.
Measure wavelength of emission lines.
Correlate lines in the hydrogen spectrum with electron transitions.
Use graphical methods to determine Rydberg's constant.

Background:
According to Bohr's model of the Hydrogen atom, the electron is found in a series of energy shells (or orbits) centered about the nucleus. Although this theory has now been superseded by the development of wave mechanics and the Schrodinger description of the electron probability density, the Bohr theory still provides an accurate means of calculating the spectrum for the Hydrogen atom.

The shells in the Bohr model are identified by the principal quantum number n, which ranges from $n=1$ (the shell closest to the nucleus) to $n=\infty$. The most stable state is where the electron is found in the lowest level ($n=1$). Through absorption of energy, it is possible to excite the electron to a higher energy level. As the electron returns to the more stable lower energy state, it loses energy in the form of electromagnetic radiation. The energy of the photon emitted corresponds to the difference in energy between the two states involved in the electronic transition. Since the energy levels of the electron are fixed, or quantized, the photon can only have a certain amount of energy. The spectrum of the Hydrogen atom will thus consist of lines at specific frequencies, rather than a broad or continuous spectrum.

Some of the photons associated with the transitions of the electron will fall in the visible region of the spectrum. This this lab, you will examine the spectrum of the Hydrogen atom and correlate the lines you observe with electronic transitions. You will use these lines to graphically determine Rydberg's constant. You will also examine the spectrum of more complex atomic gases and molecular gases. Atoms with more electrons than one will also show discrete lines in their spectra, although there will be many more lines as the energy levels become much more complicated due to the electron-electron interaction. Molecular gases (such as water vapor) will not only have electronic transitions, but also vibrational and rotational transitions in their spectrum.

If an electron is removed from the atom (ionization), the electron is free to have any energy possible and the spectrum will be continuous, rather than quantized.

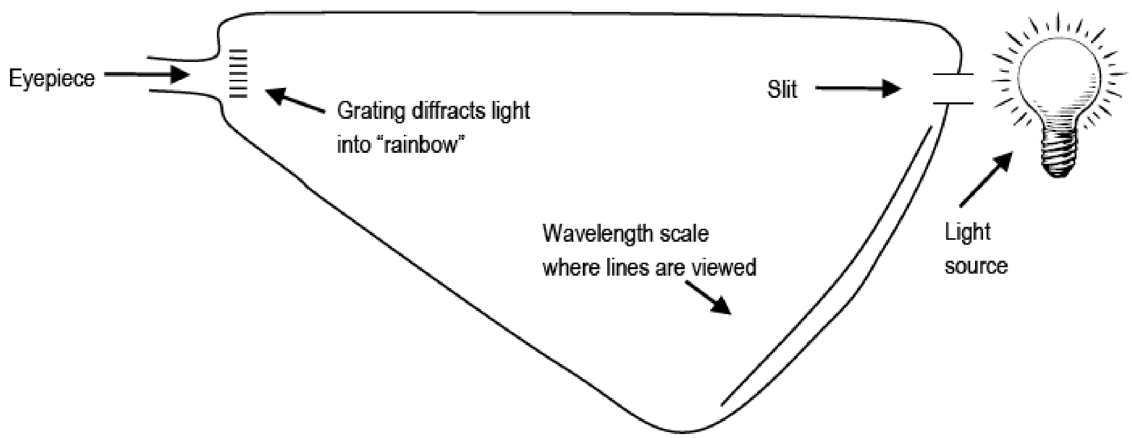
*Please read the procedure and do the pre-lab exercise before coming to lab.

Procedure:

A. Hydrogen atom spectrum.

Go to one of the lab stations where a hydrogen gas discharge tube is set up. Turn on the electricity (for no more than 30 seconds) and observe the color given off by the gas. Record your observations in the data table. If more time for observation is needed, turn off the tube for ~15 seconds and then turn on for another 30 seconds. Do not leave the tube continuously turned on; otherwise it will burn out the tube. There should be 4 lines in the visible spectrum. (You may only be able to see 3, the last blue one is difficult to see).

Hold the spectroscope so the diffraction slit is towards the light source and the diffraction grating (window) is towards you. (See diagram below).



Look through the spectroscope and observe and record the colored lines (spectrum) produced by hydrogen gas. Complete the data table below and then draw the color lines in the spectrum box that follows.

*Note: in the spectroscope the "4" refers to 400nm. The "5" to 500nm. You should estimate the wavelength to two significant figures as best you can. Record your values in *meters*. (1nm = 10^{-9}m)

*Note: these visible spectral lines represent electron transitions from energy levels 3, 4, 5 and 6 to energy level 2. For the hydrogen atom, when the electron falls from the third energy level down to the second energy level a red color is emitted. This should make sense as it is falling the shortest distance which corresponds to the lowest frequency of visible light, which is red. All other colors observed in the spectrum of hydrogen are a result of electron transitions from other energy levels (4, 5 and 6) to energy level 2.

B. Calculating Rydberg's constant for Hydrogen.

Go to an open computer and find the link provided by your instructor for the Google document (https://tinyurl.com/ydx4fls2). Plug in your 1/wavelength values (in 1/meters) in the correct location. The graph will plot $1/\lambda$ vs $1/n_i^2$. The slope of this line is $-R_H$. Record your slope.

(Why $1/\lambda$ vs $1/n_i^2$?)

The equations $\Delta E = -Rhc\ (1/n_f^2 - 1/n_i^2)$ and $\lambda = hc/\Delta E$ can be combined to give the equation:
$1/\lambda = -R(1/n_f^2 - 1/n_i^2)$

After distribution of R through the parenthesis, the following equation can be obtained which has a y=mx+b format and can thus be graphed to give a straight line.

$1/\lambda = -R(1/n_i^2) + R/n_f^2$ * rearranged the values
y = m x + b

C. Other atomic and molecular spectrum.

Using the spectroscopes, write down your observations of the other gases. Draw a picture of what each one looks like in the space provided. Note: There will be many available throughout the lab, you only need to observe 4 non-hydrogen spectrums. Make sure to record which atom or molecule you're observing.

D. Flame tests.

Obtain bottles of the following solutions:
The bottles should have a stopper with a write loop. Set up and light a Bunsen burner. Adjust the flame so that it is blue and almost invisible, rather than yellow. This is done by controlling the amount of oxygen mixing with the gas to form the flame. Your instructor will demonstrate the correct set up, technique for lighting, and adjustment of the flame for your Bunsen burner.

Label a clean test tube as "HCl" and place it in the test tube rack. Pour a small amount of HCl into the test tube. This will be used for cleaning the wire loops.

Place the wire loop into the HCl solution and then into the flame. If you see a strong color, repeat this process until the wire loop imparts little to no color to the flame.

Choose one of the solutions. Once the wire loop is clean, dip it into the solution. Make sure there is some liquid visible in the loop. Hold the loop in the flame at the top of the inner blue cone (hottest part of the flame). Record the color that you see.

You should be observing the color as the liquid on the wire loop is evaporating. If the wire loop is held in the flame too long, the metal will begin to glow and impact a color to the flame. This is not the desired color.

Repeat this process with all of the solutions listed. The color you observe is due to the cation. Write down the color of the flame for each cation.

Obtain an unknown and write down the identification letter. Perform a flame test on the unknown, and by matching the spectrum of the unknown with one of the previous salts, determine which cation your unknown contains.

Spectroscopy Experiment
Prelab Exercise

Name ____ Sec _____

✓ **Turn in at the beginning of your laboratory session.**

The value of the Rydberg constant for Helium ion (He^+) is $R_{He} = 4.39 \times 10^7\ m^{-1}$ (note this is different than the value for Hydrogen!) The energy levels of the Helium ion are given by the formula $E = -R_{He}hc(1/n^2)$ where c is the speed of light (3.00×10^8 m/s) and h is Planck's constant (6.626×10^{-34} J·s/photon).

1. Calculate the energy of the levels for n = 1 and n = 3.

2. Calculate the energy (ΔE), wavelength, and frequency of a photon associated with the transition n=3 to n=1 for the Helium ion.

3. Does the transition n=3 to n=1 involve the absorption or emission of a photon?

C. Other atomic and molecular spectrum

Draw a detailed view of the spectra (including approximate wavelengths) observed below:

Atom or Molecule: ____________________________

Atom or Molecule: ____________________________

Atom or Molecule: ____________________________

Atom or Molecule: ____________________________

D. Other atomic and molecular spectrum

Solution	**Color of flame**	**Cation Present**
NaCl	______________________	______________________
$CaCl_2$	______________________	______________________
$SrCl_2$	______________________	______________________
KCl	______________________	______________________
$BaCl_2$	______________________	______________________
$CuCl_2$	______________________	______________________
Unknown cation	______________________	______________________

Unknown Number __________

Spectroscopy Experiment:
General Questions:

1. Chlorofluorocarbons (CFCs) are of interest because of their ability to destroy ozone molecules. A carbon-chlorine bon din the CFC molecule can be broken by sunlight, leaving a highly reactive free radical which then goes on to destroy the surrounding ozone molecules. The energy of a C-Cl bond is 328 kJ/mole. Calculate the wavelength of light needed to break a bond in a single molecule. In which region of the spectrum (infrared, visible, UV) does this wavelength fall?