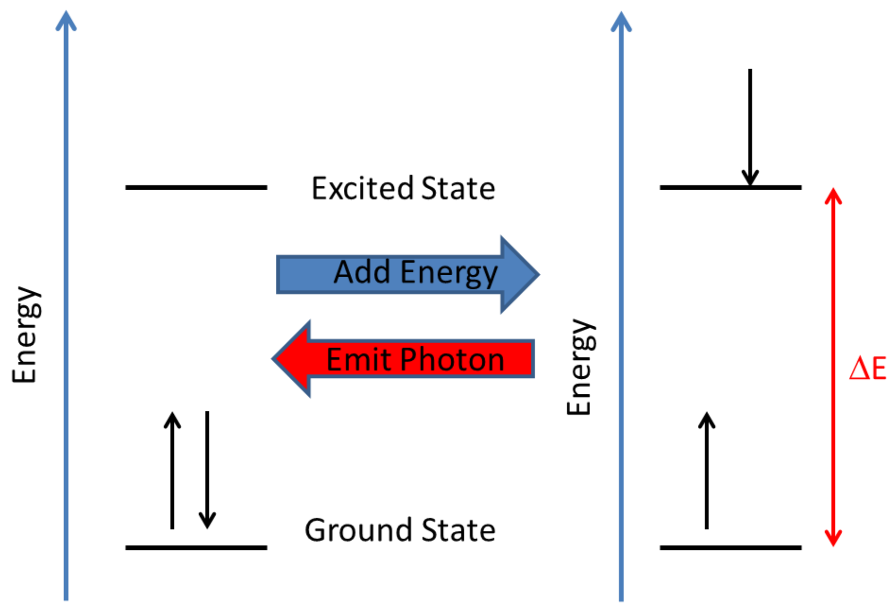
<u>Emission Spectra and the Electronic Structure of Atoms</u>

I. <u>Introduction</u>. *For background on this experiment you should read section 3.4 of your textbook, Atoms First.* In this experiment you will investigate the emission spectra from four gases. The gases are contained in a "discharge tube" through which an electric current flows. The electric current causes electrons in the gas atoms to be promoted from their low-lying "ground state" electronic energy levels to higher-lying "excited state" electronic energy levels.

An electron in an excited state will return to the lower-lying ground state. When the electron returns to the ground state it will release energy either in the form of heat or in the form of light. In the case of the gases you are studying, the relaxation of an electron from the excited state back to the ground state results in the release of light. Each electron that undergoes a transition from the excited to the ground state will release energy in the form of a particle of light called a photon.



Simple two-level energy diagram. On the left, two energy levels are shown with two electrons occupying the ground state energy level. Upon the addition of energy equivalent to ΔE, an electron is promoted to the higher-lying energy level called the excited state as shown on the right. The electron in the excited state will spontaneously, return to the ground state, releasing a photon with energy ΔE.

The figure above shows two configurations. On the left there are two electrons of opposite spin, represented by arrows, which occupy the lower energy level. This is the ground-state configuration for this system. Addition of energy in the form of heat, electrical work or light is required to promote one of the electrons to the excited state energy level. The configuration on the right is the excited state configuration. The electron in the excited state is not stable. Therefore, it will return to the lower energy level, much like a ball will roll to the bottom of a hill. One way for it to return to the lower level is to release energy in the form of a photon. Photons have discrete energies given by the following equation.

$$E_{photon} = h\nu = h\frac{c}{\lambda}$$

In this equation, h is Planck's constant (6.63×10^{-34} J s), ν is the photon frequency in Hertz (s^{-1}), c is the speed of light (2.99×10^{8} m/s) and λ is the wavelength of light associated with the photon energy.

If a photon is released when an electron returns to a lower energy level, the energy of the photon will match the energy difference of the two levels.

$$E_{photon} = \Delta E_{electron} = E_{excited\ state} - E_{ground\ state} = h\nu = h\frac{c}{\lambda}$$

You will be observing the line spectra from electronic transitions in gas phase atoms (Ne, He, Hg and H_2). In these atoms, electrons are promoted by electrical current and when they return to their ground state they emit a photon. As you have learned in class, an atom is not a two-level system like the simple example above, but rather a series of electronic energy levels or "shells". Each shell has a quantum number, *n*, associated with it that determines the energy of an electron in that shell. For the atoms in the gases you are studying, there are many possible transitions between different shells. Each emissive transition will involve an initial excited state (higher level shell with quantum number n_i) and a final ground state (quantum number n_f). For emission, the value of n_i is always greater than n_f. Each transition leads to an emission line in the line spectrum. Each line spectrum is composed of lines with wavelengths that follow the Rydberg formula (equation 3.9 in your textbook).

$$\Delta E = h\nu = \frac{hc}{\lambda} = R_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \quad \text{Rydberg Formula}$$

In the Rydberg formula, the quantities n_f and n_i are the principle quantum numbers of the final and initial states of an electronic transition. The quantum numbers are always positive integers with the lowest possible value of 1. In this formula R_H is the Rydberg constant and you will estimate its value from the spectrum that you measure for hydrogen.

In this experiment, you will use a spectroscope which is an optical device that contains a slit through which light passes. The light impinges on a diffraction grating that separates the light into its different wavelengths. This results in a spectrum much like what you see when light passes through a prism or when you see a rainbow. This spectrum strikes a screen with a scale on it. On the spectroscope there is an eye-piece through which you can view the screen and scale. The spectrum will appear as colored vertical lines appearing on the scale.

Color	**Approximate Wavelength Range**
Red	630 nm and greater
Orange	590 to 630 nm
Yellow	560 to 590 nm
Green	510-560 nm
Bluish-green(Teal)	470-510
Blue	400-470

To use the spectroscope, point the entrance slit at the light source and look through the eye piece. You should see a spectrum in the form of colored bars or lines projected onto the scale. If you are looking at white-light from an incandescent bulb or from reflected sunlight you will see a continuous spectrum of blue to red. If you look at a low-pressure discharge lamp that your TA has set up, then you will see discreet lines from the atomic emission of the gas contained inside. To practice using your spectroscope, aim it at the fluorescent lamps in the lab and look through the eye-piece. You will observe a line spectrum that comes from the lamps.

II. <u>Procedure:</u> Your TA will set up discharge tubes and you will view the tubes through your spectroscope. Record the number on the scale for each line you see and record the color of the line. To help you make line assignments on your data sheet it helps for the scale to be calibrated with respect to color and wavelength. Most people see roughly 400 nm to 700 nm light, the visible spectrum. Familiarize yourself with the wavelength ranges of different colors in the table above.

Name ______________________________ TA ______________________________ Time ____________

Reading	He λ (nm)
	439 w
	444 w
	447
	471 w
	492
	504 w
	587
	667

Examine the photons emitted from all four of the lamps provided. Take a reading from the scale for each line observed in the spectra projected on the scale. Try to match the observed lines to each of the wavelengths listed in the three boxes on the left which came from a reference book. ("w" means a weak intensity line that may not be seen). Record the reading in the corresponding table. For example, the scale readings for the Hg spectrum on the last page are different from the results that you will have obtained.

Reading	Hg λ (nm)
	405
	408
	436
	492 w
	546
	578
	690 w

Then record the four readings from the scale for the photons emitted from hydrogen and fill them in the box to the right.

reading	**H λ (nm)**
	To be calculated
	To be calculated
	To be calculated
	To be calculated

Reading	Ne λ (nm)
	540
	585
	588
	603
	606
	616
	621
	626
	633
	638
	640
	651
	660
	693
	703

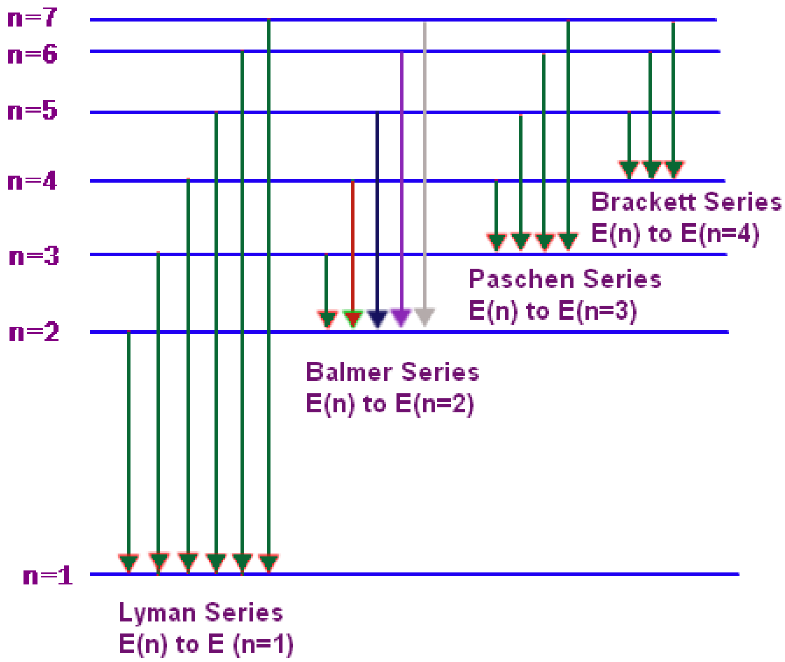
Construct a graph using Excel for each of the three elements you are evaluating for a standard; He, Hg, and Ne. (an example is on the last page)

1) Use the wavelength as the y axis and the reading from your scale as the x.

2) Find the equation of the trend line and R^2. The R^2 closest to 1 is your best fit line, use whichever element has the best R^2 as your standard.

Print your calculations and all three graphs and attach them to your lab worksheet along with all your calculation pages. The way to do this is to neatly arrange all three graphs and your calculations in one area. Then highlight all the squares with your calculations and the area behind three graphs. Then choose print from the pull down menu.

Now using your BEST graph as the standard, use the forecast function in excel to calculate the wavelengths for the hydrogen atom from your readings. (The instruction is on the last page)



$$\Delta E = h\nu = \frac{hc}{\lambda} = R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

Rydberg formula

n_f = Final State

n_i = Initial State

Data table for hydrogen

Spectroscope Scale reading	λ calculated	λ actual (from literature)	% error	Photon Energy (cm^{-1})*	n_f	n_i	$(\frac{1}{n_f^2} - \frac{1}{n_i^2})$
		656.2			2		
		486.1			2		
		434.0			2		
		410.1			2		

- To convert from nm to cm^{-1}, take the reciprocal of the wavelength and multiply by 10^7.

The line spectrum you are observing at visible wavelengths for hydrogen are part of the Balmer Series (see diagram on next page). For the Balmer series the quantum number of the final state (ground state) is 2, i.e. n_f = 2. The lines you observe come from transitions from excited states that have quantum n_i = 3, *4*, *5*, 6. Obtain the actual wavelengths and calculate the percent error for each line in the table above. Calculate the photon-energy of each of the four lines observed, fill in the values of n_f and n_i describing the energy transition which produced the line and calculate the function $(\frac{1}{n_f^2} - \frac{1}{n_i^2})$.

Now using excel, plot the data in the $(\frac{1}{n_f^2} - \frac{1}{n_i^2})$ column vs. the Photon Energy in cm^{-1} to obtain R_H from the slope. Attach your Excel graph to the report.

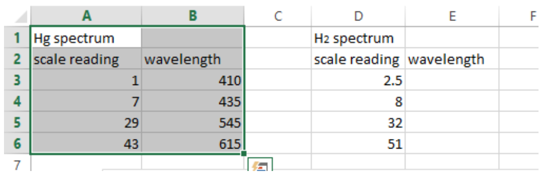
R_H = ___________________

Using Excel

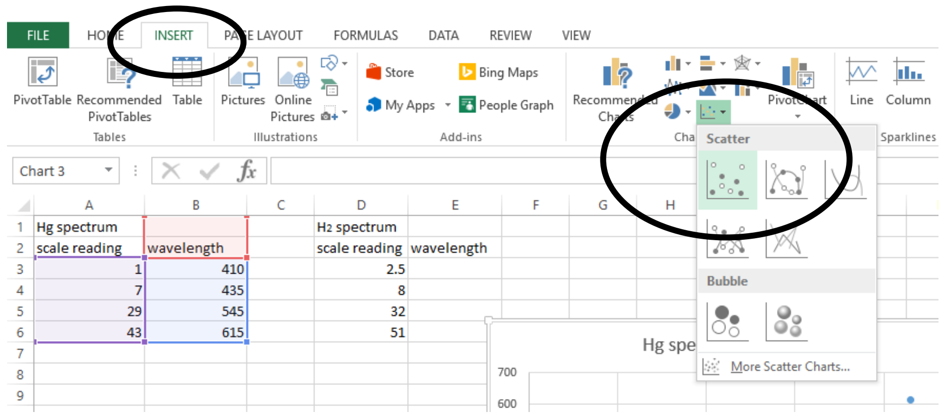
1. Making graph from data that you obtained.
   a. Having data put in Excel spreadsheet as table 1. Hg Spectrum.
   b. You will need two more tables, one for He Spectrum, and one for Neon Spectrum.
   c. Select entire data table that you need to graph

Hg spectrum	
scale reading	wavelength
1	410
7	435
29	545
43	615

Table 1. Hg Spectrum

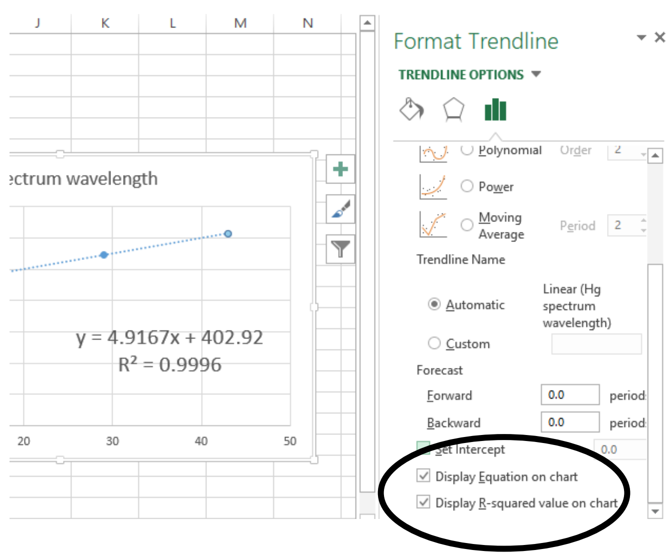


   d. Select from tab menu Insert > Charts > Scatter chart



   e. Right click on any data point on your graph, select Add trendline from scroll down menu.

   f. On the right side, pick option of display equation, and display R^2 value.
   g. Ideally, R^2 is 1. Thus, your BEST graph is the one has R^2 value is closest to 1.



2. Finding Hydrogen wavelengths (λ calculated) using Forecast function in Excel.
   a. Having Hydrogen spectrum data as table 2.
   b. To find corresponding wavelength to the scale reading, you need to type in "= forecast (x, known ys, known xs)"
   c. Where x is your corresponding scale reading.
      Known ys are all y values from your best graph
      Known xs are all x values from your best graph

H_2 spectrum	
scale reading	wavelength
2.5	
8.0	
32.0	
51.0	

Table 2. Hydrogen Spectrum