

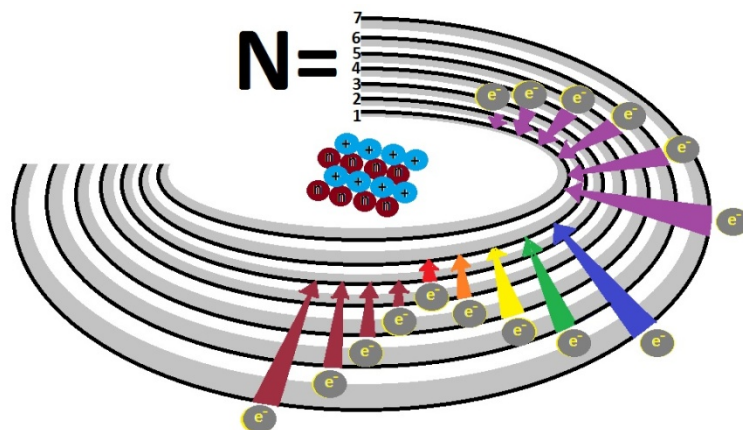
<b>Atomic Emission Spectroscopy</b>	<b>Name:</b>	
	<b>Lab Section:</b>	
	<b>Date:</b>	<b>Sign-Off:</b>

by

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## Atomic Emission Spectroscopy

Atomic spectroscopy is the study of the electromagnetic radiation that is absorbed and given off by electrons. Electrons can sometimes become “excited” (i.e. the atom absorbs energy, and the electrons move up at least one energy shell). Because we know that what goes up must come down (Think, *potential energy*. Not the formula, just the concept.), we know that the electron that has gone UP an energy level, must also come DOWN an energy level. What does this mean? It means that just as the electron absorbed energy, it will also give off this energy to its surroundings, but this time it will be observed by you as visible light!

Each atom has its own unique color spectrum that it gives off, and can be used to identify said atom. The gas-discharge tubes you will be using today, work by exciting the atoms with an electrical current, which will then (as read above) allow the electrons to drop from their higher energy orbits to lower orbits, giving off light.

It is important to note that different electrons will move different distances to and from the nucleus, which means lower and higher energies, respectively. Knowing then that the energy is dependent upon the distance from the nucleus, we can conclude that if electrons are moving different distances, that each atom can give off multiple wavelengths of light radiation. Using this information, we must now figure out (quantitatively) what kind of energy is being given off by each atom.

The Rydberg equation for the energy of a specific wavelength of light ( $E_n$ ) is as follows:

$$E_n = -R(1/n^2)$$

Where  $\beta = 2.2 \times 10^{-18} \text{ J}$  and,  $\nu = c/\lambda$ ,  $\nu$  = frequency,  $\lambda$ =wavelength and  $c$ =speed of light  $= 3.00 \times 10^8 \text{ m/sec}$ . Planck's constant,  $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{sec}$ , is also taken into account with Rydberg's equation, as well as light. During calculations, the final wavelength will be taken as an absolute value (  $|x|$  ) because the Rydberg equation is a prediction of the amount of the energy that will be given off between two states (i.e. light given off is positive energy). We also know that the Balmer-Rydberg equation,

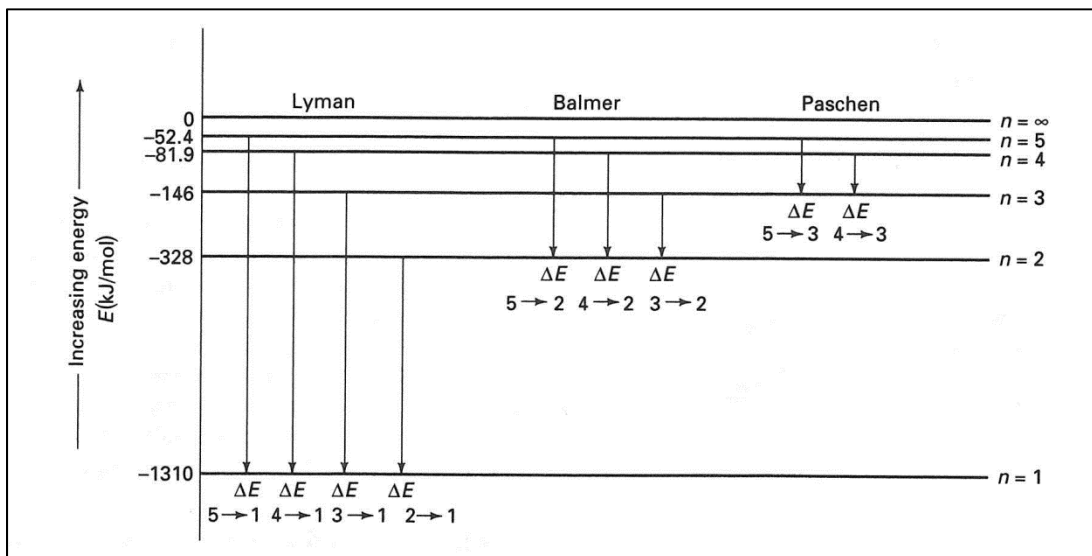
$$1/\lambda = R (1/n_f^2 - 1/n_i^2),$$

will allow us to determine the wavelength of light associated with specific energy transitions. This is important, because that will allow us to determine the energies of electrons as they pass through their energy levels to give off light radiation. It is important to note, above, that  $n_i^2$  = **initial energy level**,  $n_f^2$  = **final energy level** and **Rydberg constant =  $R = 1.097 \times 10^{-2} \text{ nm}^{-1}$** .

Using Figure 1 (below), and determining the **E** for each wavelength (also found below), you will be able to get a rough estimate of where in the Balmer series the different wavelengths occur. This will help you understand why the different wavelengths give off color the way that they do. Towards the left of the Balmer series the wavelengths are shorter, which means that as the wavelengths approach Ultra-Violet, more energy given off. As you move towards the right of the Balmer series, the wavelengths become longer as they approach Infra-Red, which means that they give off less energy. Here are some approximations for the Balmer series, to see where your wavelengths are located:

1.  $n=5 \rightarrow n=2$  – Energy:  $4.58 \times 10^{-19} \text{ J}$  Wavelength:  $4.34 \times 10^{-7} \text{ m}$
2.  $n=4 \rightarrow n=2$  – Energy:  $4.09 \times 10^{-19} \text{ J}$  Wavelength:  $4.86 \times 10^{-7} \text{ m}$

3.  $n=3 \rightarrow n=2$  – Energy:  $3.03 \times 10^{-19} \text{ J}$  Wavelength:  $6.56 \times 10^{-7} \text{ m}$



(Postma, J M, J L Roberts, J L Hollenberg. *Chemistry in the Laboratory 6<sup>th</sup> ed.* 2004. WH Freeman.)

**Figure 1**

We also know that light can come in two distinct forms, which means that light has “dualistic properties.” (Search for “Young’s Two-Slit Experiment” for better understanding.)

What we have gone over so far is the wave property of light, and now we will briefly discuss the photon property of light. Photons are like little “packets” of light, if you will, that give off energy. This change in photon energy can be calculated with the equation:

$$\Delta E = h \nu = hc / \lambda = R (1/n_f^2 - 1/n_i^2)$$

Remember that E (energy) can be calculated using Einstein’s equation,  $E = mc^2$ , as well as with Planck’s equation,  $E = h\nu$ . Eventually, this translates to  $mc^2 = h\nu$ .

Using the spectroscope that each group is given, you will be determining, by group, the wavelengths of light present, and then determining the change in energy that the electrons give off. The inside of the spectroscope has been labeled with the approximate wavelength of light with which you will record your data.

### **Experimental**

Using your ring stand base and rod, attach a three-pronged clamp about 20 cm above the base (adjust per individual height difference). Place the base of the stand approx. 25-30 cm from the lamp, and then securely fasten (without cracking) the spectroscope to the three-pronged clamp.

**Follow these instructions to the letter!**

**DANGER: High Voltage! Keep your fingers out of the discharge tube sockets!**

To use the power source with a gas discharge tube:

1. Make sure that the power source is off before plugging into the outlet.
2. Carefully unpack the gas discharge tube and insert it into the bottom end of the power supply (socket) and press down carefully.
3. Press back away from yourself as you are pressing down and let the top of the discharge tube lift into the upper socket in the power supply.

4. At this point, you will want to get someone with good eyes so that you can line up the spectroscope with the light source before turning the light source on.
5. Before you begin looking at the light, make sure to put on your UV protection goggles which will be provided by your lab instructor to the individual looking through the spectroscope.
6. Plug the power source into the outlet.
7. Turn the power source on for NO MORE THAN 30 SECONDS.
8. Once 30 seconds have been reached, turn the power source off for a MINIMUM OF 30 SECONDS.
9. You may then turn the power source on again as needed, provided you follow the 30:30 rule.
10. Turn the power source off.
11. Unplug the power source from the outlet once the power has been turned off.
12. Remove the gas discharge tube, reversing insertion steps.

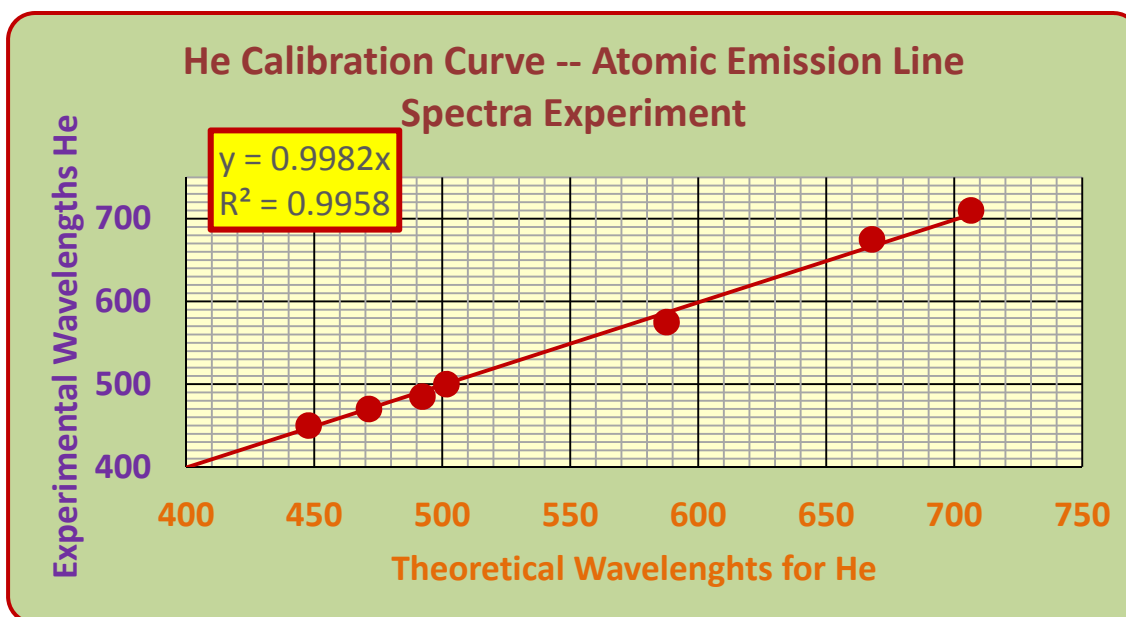
These rules are for your safety, as well as for the longevity of the equipment.

To use the spectroscope, you will look through the eyepiece (narrow side) so that you are able to read the colored vertical lines beneath the labeling on the inside. Record these values to three significant figures, in the table, below. Remember that once you have aligned your spectroscope with your light source, it is important to not move it. The viewer will simply verbalize the color that is visible, followed by the approximate reading

of the wavelength. You may wish to do several trial runs so that you can attain the most accurate readings.

You will calibrate your spectroscope by using the He tube and making a scatter-plot in Microsoft Excel that contains a best-fit trend line between your experimental and theoretical data points. This line will then be used to determine the corrected wavelengths of the other gasses. Record your data in the table below.

Color	Actual Wavelength (nm)	Actual Wavelength (Å)	Observed/Experimental Value (from spectroscope scale)
Violet	447.76	4478	
Blue	471.314	4713	
Blue	492.193	4922	
Green	501.567	5016	
Yellow	587.562	5875	
Orange/red	667.815	6678	
Red	706.5	7065	



**Figure 2.** Calibration Curve for Spectroscope Using He.

Using the above method, use your spectroscope to determine the values of the following gases: Group 1, Hydrogen; Group 2, Neon; Group 3, Argon; Group 4, Mercury. Then, using your calibration curve (e.g., **Figure 2**, bottom, previous page), you will use a straight-edge (i.e. a ruler) to draw a line from your experimental axis to your best-fit line and then over to your theoretical axis for the corrected value (or, better, yet, use the line equation that Excel provides to calculate the corrected value). Once you have determined the corrected values, calculate the percent error for your readings.

**Percent error = (experimental value – theoretical value) / theoretical value x 100**

### Hydrogen spectrum

Color	Scale reading	Wavelength (nm) from calibration graph	Known wavelength (nm)	% error
Red			656.2	
Turquoise			486.1	
Blue/Violet			434.0	
Purple			410.1	

### Neon Spectrum

Color	Scale reading	Wavelength (nm) from calibration graph	Known wavelength (nm)	% error
Dark Red			659.9	
Red			640.2	
Orange			607.4	
Yellow			588.2	
Yellow			585.2	
Green			540.1	

### Argon Spectrum

Color	Scale reading	Wavelength (nm) from calibration graph	Known wavelength (nm)	% error
Green			514.5	
Aqua			501.7	
Light Blue			488.0	
Blue			476.5	
Dark Blue			457.9	



## Mercury Spectrum

Color	Scale reading	Wavelength (nm) from calibration graph	Known wavelength (nm)	% error
Violet			435.8	
Green			546.1	
Yellow			577.0	
Yellow			579.1	

Now you will calculate the photon energy for each wavelength your group measured, to determine the initial and final electron states for each spectral line found in the spectroscope. You will determine the photon energy in **kJ/mol**, so don't forget about Avogadro's number! Let's use Neon's line at 539 nm for our example calculation of photon energy work:

Neon – 539 nm

$\nu = c/\lambda$  and  $\nu = \text{Speed of light in a vacuum (c) / wavelength of light } (\lambda)$

$$\nu = 3.00 \times 10^8 \text{ m/sec} / 5.39 \times 10^{-7} \text{ m} = \underline{5.5659 \times 10^{14} \text{ s}^{-1}} = \nu$$

$$E = h\nu \rightarrow 6.626 \times 10^{-34} \text{ J}\cdot\text{sec} (5.5659 \times 10^{14} \text{ s}^{-1}) = \underline{3.688 \times 10^{-19} \text{ J}} = E$$

$$3.688 \times 10^{-19} \text{ J} (6.023 \times 10^{23}) = 222124.6642 \text{ J/mol}$$

$$222124.6642 \text{ J/mol} (1\text{kJ}/1000\text{J}) = \mathbf{222.124 \text{ kJ/mol}}$$

Now let's determine the transition ( $n_i$ ) using the known value of R ( $1.097 \cdot 10^{-7} \text{ m}^{-1}$ ):

$$\frac{1}{\lambda} = R \left( \left( \frac{1}{n_f^2} \right) - \left( \frac{1}{n_i^2} \right) \right)$$

$$\frac{1}{5.39 \cdot 10^{-7} \text{ m}} = (1.097 \cdot 10^7 \text{ m}^{-1}) \left( \left( \frac{1}{2_f^2} \right) - \left( \frac{1}{n_i^2} \right) \right)$$

$$1855287.57 = (1.097 \cdot 10^7 \text{ m}^{-1}) \left( \left( \frac{1}{4} \right) - \left( \frac{1}{n_i^2} \right) \right)$$

$$\frac{1855287.57}{1.097 \cdot 10^7} = 0.25 - \frac{1}{n_i^2}$$

$$\frac{1855287.57}{1.097 \cdot 10^7} - 0.25 = -\frac{1}{n_i^2}$$

$$-0.080876 = -\frac{1}{n_i^2}$$

$$n_i^2 = \frac{1}{0.080876} = 12.365$$

$$n_i = 3.52 \rightarrow 4$$

Now that we know the initial transition state (which we could have intuited based on the color of the light), let's use the calculated energy (E) and the calculated  $n_i$  to determine R:

$$E = R \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

$$222,124 \frac{\text{J}}{\text{mol}} = R \left( \frac{1}{2_f^2} - \frac{1}{4_i^2} \right)$$

$$222,124 \frac{\text{J}}{\text{mol}} = R (0.25 - 0.0625)$$

$$\frac{222,124}{0.1875} = R = 1.185 \cdot 10^6 \text{ m}^{-1}$$

## Results

Tabulate your calculations for the gas that your group studied in the table below:

Gas Discharge Tube Used:				
Color	Corrected Wavelength	$n_i$	Energy (J/mol)	R Calc'ed

Confer with the other 3 groups and record their data in the tables, below.

Gas Discharge Tube Used:				
Color	Corrected Wavelength	$n_i$	Energy (J/mol)	R Calc'ed

Gas Discharge Tube Used:				
Color	Corrected Wavelength	$n_i$	Energy (J/mol)	R Calc'ed

Gas Discharge Tube Used:				
Color	Corrected Wavelength	$n_i$	Energy (J/mol)	R Calc'ed

Do NOT include the He data that each group should have obtained on their own.

For each gas, determine the average R:

Gas: \_\_\_\_\_ Average R: \_\_\_\_\_

Gas: \_\_\_\_\_ Average R: \_\_\_\_\_

Gas: \_\_\_\_\_ Average R: \_\_\_\_\_

Gas: \_\_\_\_\_ Average R: \_\_\_\_\_

## Questions

- 1) Explain briefly why  $R$  for Neon at 539 nm is an order of magnitude greater than the theoretical value.
- 2) Explain why for this experiment that  $n_f$  is 2.
- 3) If Einstein's and Planck's equations equate, what does that mean?
- 4) Explain briefly and succinctly what the first quantum number means to a chemist when dealing with atoms.

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