CHEMISTRY 101L DATA SUMMARY

EXPT. 8

Spectroscopy

Your name:	Joe Puccio	Lab Section	425
Your partner's name	Jake Evans	Date	11/11/15

Results

Table 1. Fluorescent Light Position on Spectroscope by Wavelength

Color of Line	Position on Spectroscope	Wavelength (nm)
Blue	5.05	435.8
Green	5.60	546.1
Yellow	6.05	578
Equation for Calibration Line	Position on Spectroscope (no units) = 0.0065 * (wavelength in nm) + 2.1855	

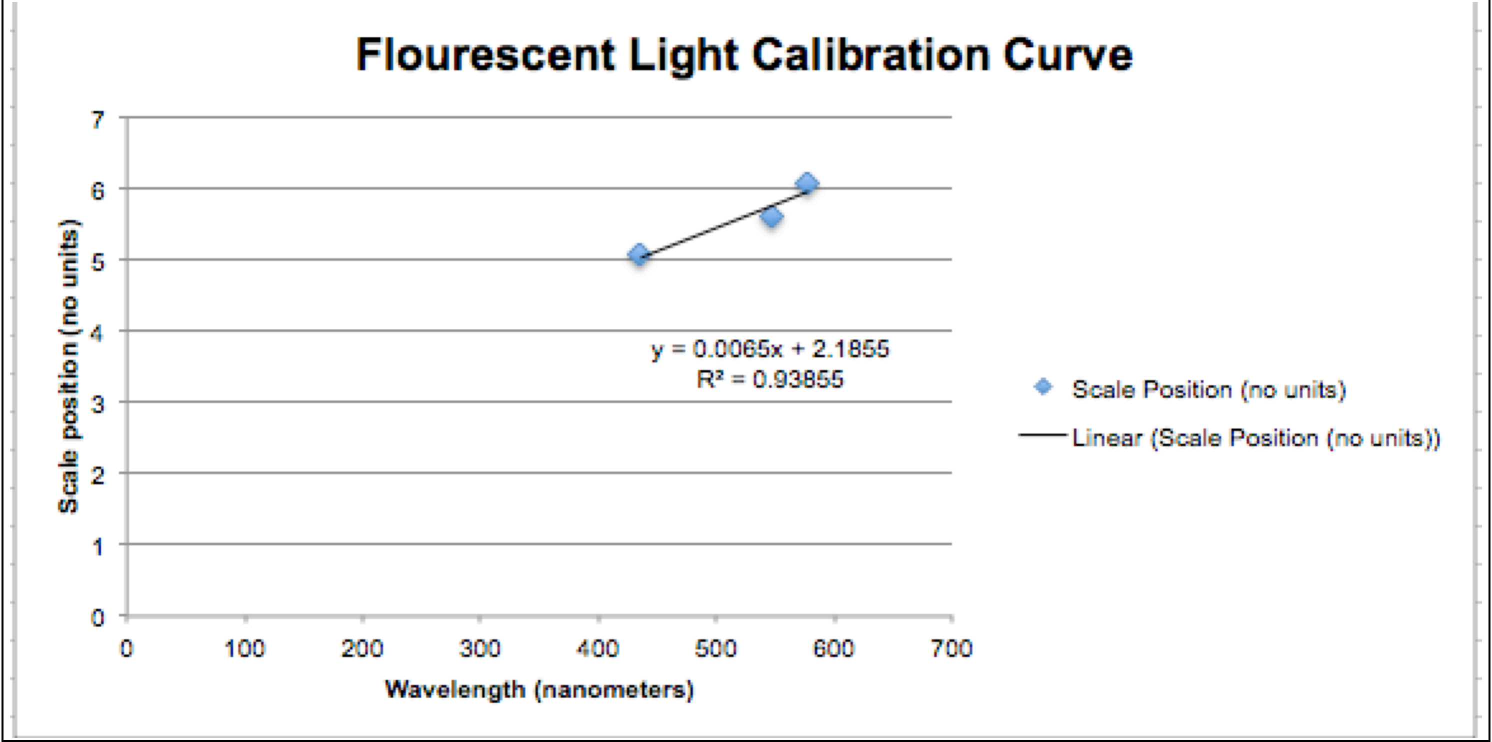

Figure 1. A plot of the scale position of constituent frequencies making up fluorescent light plotted as a function of wavelength.

Table 2. Wavelength range observed emitted from variety of light sources and the wavelength regions with maximum intensity.

Light Source	**Wavelength Range (nm)**	**Region(s) of Maximum Intensity (nm)**
Candle	758 - 326	758-619, 588-465
Lightbulb	780 – 326	uniform
Hydrogen	719 - 349	uniform

Calculation: By taking the functional inverse of the equation for the calibration line, we are able to compute the wavelength range provided a position on our spectroscope. For instance, provided a spectroscope position of 6.0, we may calculate the wavelength = (6.0 - 2.1855)/0.0065 = 588 nanometers.

Table 3. Spectroscope position observed for Hydrogen emission and the calculated and literature wavelengths, along with calculated percent error. Literature wavelengths were retrieved from: http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch6/bohr.html on 11/11/15.

Color of Line	**Position on Spectroscope**	**Calculated Wavelength (nm)**	**Literature Wavelength (nm)**	**Percent Error**
red	6.85	719	656	-9.6%
blue	5.05	442	434	-1.8%
violet	4.45	349	410	14.8%

Calculation: We must calculate the expected (literature) minus the calculated, divided by the expected (literature). So, for example, (434 – 442) / 442 = -.018 = -1.8%.

Table 4. Frequency, energy, and transition results based on observed wavelengths for Hydrogen omission.

Color of Line	**Wavelength (nm)**	**Frequency (Hz)**	**Energy (J)**	**Transition**
red	719	4.17 * 10^14	2.765×10^-19	3 -> 2
blue	442	6.783 * 10^14	4.497×10^-19	5 -> 2
violet	349	8.59 * 10^14	5.695×10^-19	6 -> 2

Calculation: Frequency is calculated by taking (speed of light in m/s) / 442nm = 6.783*10^14 Hz. The energy of that frequency can be calculated by taking Planck's constant (6.63 * 10^-34 J*s) multiplied by the frequency, so (6.63 * 10^-34 J*s) * 6.783*10^14 Hz = 4.497×10^-19 joules. We can use 2.18*10^-18 * (1/n^2_final – 1/ n^2_initial) to calculate the expected energy release for two states, for instance we see 2.18*10^-18 * (1/2^2 – 1/5^2) = 4.6 * 10^-19 Joules, which is approximately our discovered energy for the blue line, which indicates that there was a 5-> 2 transition.

Table 5. Observations and calculated Wavelength of various known elements subject to Bunsen burner

Element	**Color of Flame**	**Color of Emission Lines**	**Position**	**Wavelength**
Iron	Yellow	Yellow/Blue	6.1/5.3	603/480
Lithium	Red	Red/Yellow/Blue	7/6.1/5.35	742/603/488
Sodium	Yellow/Orange	Yellow	6.15	611

Copper	Green	Yellow/Green/Violet	6.1/5.5/4.5	603/511/357
Potassium	Violet	Red/Yellow/Green/Violet	6.4/6.1/5.6/4.35	649/603/526/334

Table 6. Observations and Wavelength calculation for unknown mixture in Bunsen burner

Color of Flame	**Color of Emission Lines**	**Position**	**Wavelength**	**Elements in Unknown Mixture**
Yellow	Yellow/Green/Violet	6.1/5.7/4.5	603/542/357	Copper

The color of light corresponds to the frequency (and therefore energy) of the photons that makeup that light. Higher frequencies, thus higher energy, light corresponds to blue/violet appearances, while lower energy light corresponds to orange/red appearance. When an electron in an atom is excited by a current or absorption, some of its electrons may be raised to a higher energy state (raising their distance from the nucleus, essentially moving up a shell), and when these atoms leave the higher energy state and the electrons move down to the lower energy state, a photon is emitted that corresponds to the state change of the electron, which is also tied to the type of atom that was excited. That is, exciting a Copper atom may release different energy photons than an Iron atom.

It's possible that some groups only saw three lines in the visible spectrum of the hydrogen emission lamp rather than four because the last line is very difficult to discern, as it's nearing the edge of the visible spectrum. It's also possible that the lamp was not able to excite the electrons to the necessary degree to omit that frequency of light, however the former explanation is more likely.

Because Neon has 10 electrons, it has two shells full completely, while hydrogen has only one electron which means its innermost shell is not yet full, and thus this electron is orbiting much closer to the positive nucleus compared to neon. Thus, when excited, neon will have a broader spectrum as it has many more electrons at very different energy levels that could be excited and thus emit a larger range of wavelengths, while hydrogen only has one electron that could be excited and thus very few potential state changes.

In addition to the color of the flame, it was clear that Copper was the makeup of the unknown mixture based on the fact that the spectrum of the unknown mixture fully matched the spectrum of Copper, down to the wavelength with very little difference. However, it's possible that there was another element present which had wavelengths that were a subset of those observed in the unknown mixture, such as sodium. This is because it's difficult to spectrally discern if sodium was present in addition to Copper as they both emit light at wavelength ~611nm. The other elements could be ruled out because they all produced at least one wavelength that was not observed in the unknown mixture's spectrum.

Because many individuals perceive color differently, it's possible that we misidentified the color of certain flames. It's also possible that we did not correctly calibrate our spectrometer due to variation in the florescent lighting from the expected spectrum produced by florescent lighting, this would lead us to invalid wavelength calculations for nearly all of our wavelength values. Some natural light could have affected this calibration, as the light source during calibration was intended only to be florescent. During measurement, it's also possible that natural light was present in the room and thus included in the recorded spectra for various measurements. It's also very possible that our spectrometer value ranges were incorrectly identified, as we did achieve some wavelengths that should not have been visible to the human eye (ultraviolet/infrared).