CHEMISTRY 101L DATA SUMMARY

EXPT. 7

Beer's Law

Your name:	Joe Puccio	Lab Section	425
Your partner's name	Jake Evans	Date	11/03/15

Results

Record the identity of the herb you tested: Oregano

Observations: There was a very linear gradient of transparency formed by the standard solutions in the cuvettes. An increase in the concentration of $FeCl_3$ in the solution corresponded to a decrease of transparency and an increase of a reddish/orange hue. Interestingly, our final solution prepared with the herb, fell visually on this spectrum at around the transparency of our S2 sample. This indicates that the two may have similar concentrations of $FeCl_3$. There was some solid remaining in the filter paper after pouring the pyrolyzed herb sample, but there was no solid in the crucible, as all of the solid was filtered out successfully by the paper. It's difficult to tell if there actually were small solid pieces in the solution, however.

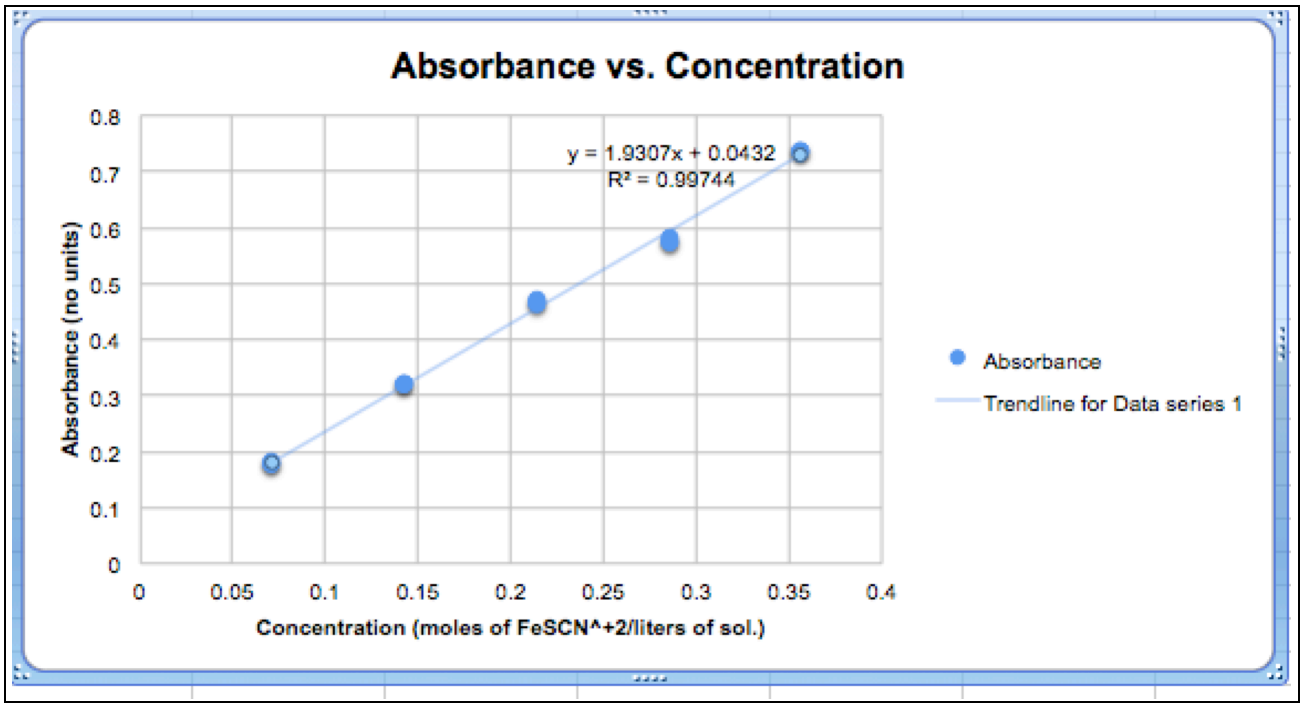

Figure 1. Absorbance vs Concentration of our samples. We can see that as concentration of $FeSCN^{2+}$ increases, so does the light absorbance of the solution. As expected from the literature, the relationship between concentration and absorbance is linearly positively related. With an R^2 of .997 > .95, we can conclude that our data is a good linear fit. A literature value of the molarity coefficient of $FeSCN^{2+}$ obtained from Collin.edu is 7260 M^{-1} cm^{-1}.

Table 1. Calculation of experimental molar extinction coefficient.

Mass of herb sample (g)	.20 grams				
%T readings	57.3	57.4	57.4		
Absorbance values (no units)	0.242	0.241	0.241		
Calculated Solution Concentration (mM)	0.103	0.102	0.102		
Mean Concentration ± standard deviation (mM)	0.10233 +/- 0.00058				
millimoles Fe^{3+}	.002575				
mg Fe^{3+}/100 g	71.9mg				
USDA mg Fe^{3+}/100 g	36.72mg				
% error	95.8%				

We calculate the Absorbance from the %T readings by calculating: 2-LOG(%T,10), which yields a unit less value of absorbance. Then, using this value for absorbance, we calculate concentration based on the line of best fit we computed. Namely, subtracting the y intercept and dividing by the slope to yield the concentration: (absorbance – 0.0432)/1.9307, which yields the concentration. So, for instance in the table above we have absorbance = 2-log(57.3,10) = 0.242 and concentration = (0.242-0.432)/1.9307 = 0.102 (mM). We had a solution of 25mL of solution, which at .102mM concentration means we had .002575 millimoles of Fe^{3+}. That's the number of millimoles for .2 grams of herb, so multiplying it by 500 gives the number of mg per 100g of herb, which is 71.9mg.

Conclusion:

By experimentally determining the relationship between concentration of iron in a solution and the absorbance of that solution to certain wavelengths of light, we were able to approximate the concentration (and consequently, the amount as the volume was known) of iron in a substance with an unknown concentration. By Beer's law we know that the relationship between concentration and absorbance is linear, which enabled us to be confident in a linear fit of our experimental data. We then used this linear fit to predict the concentration of the solution based on its absorbance. We experimentally concluded that there were 71.9mg of iron in a sample of 100g of Oregano (extrapolated from conducting on a sample of .2 grams). We took three measurements of absorption to assist the chance that our measured values were not due to some random, confounding variable.

The $FeCI_3$ is the limiting reagent. We know that this is the case because the extent to which the solution turned dark red was a function of the $FeCI_3$ that we added to each of the cuvettes, not the amount of the other reagent. So long as the total volume of solution was constant across each of the cuvettes, then adding more $FeCI_3$ to the solution means an increase in the concentration of $FeCI_3$. The setup depends on this being true because we are relying on the fact that a higher amount of $FeCI_3$ will result in a darker (red) sample, due to the increase in the dark red reactant $FeSCN^{2+}$. If $FeCI_3$ were not the limiting reagent, we would not have seen a correlation between absorption (due to the solution being darker) and $FeCI_3$ concentration. Thus, we must conclude that it is the amount of $FeCI_3$ in the solution that dictates the degree to which the solution turns red. It is because $FeCI_3$ is a limiting reagent that an excess of the other reagent does not change the results.

Error could have been introduced while using the Bunsen burner, as some of the iron could have been lost due to small particles of the herb being ejected into the air. This would have resulted in a lower than

expected amount of iron. Another possible source of error is that our cuvette absorption measurement values could have been off due to a contaminated sample of $FeCI_3$, as many students could have accidentally contaminated this sample. If the sample were less concentrated than the advertised molarity due to contamination, then we would have achieved an estimated iron concentration that is too high, as we are overcompensating. Moreover, we ran through the experiment twice due to a small measuring error, and so improperly cleaning the cuvettes (leaving some residue) could have increased the measured iron concentration.