

## EXPERIMENT B7: GREEN CRYSTAL

### Learning Outcomes

Upon completion of this lab, the student will be able to:

- 1) Synthesize an inorganic salt and purify it using the method of recrystallization.
- 2) Test the purity of the synthesized crystals.
- 3) Analyze the synthesized crystals and predict its formula using methods such as: titration, spectrophotometry, and determination of percent hydrate.

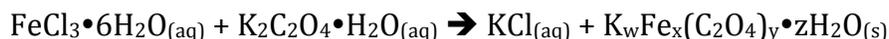
### Introduction

One of the most exciting aspects of chemistry is the ability to synthesize new materials. These synthetic materials are in a variety of consumer products and can be found in products such as food, health and beauty supplies, and pharmaceuticals. Chemists employ conceptual chemical models to formulate relatively simple procedures to synthesize such substances reliably and in high yield. Once synthesized, the substances must be purified and analyzed to confirm the chemical identity. This ensures the quality and effectiveness of the product in question. Accordingly, synthesis, purification, and analysis are three essential methodologies every chemist must be familiar with.

In this experiment a Green Crystal will be synthesized from iron (III) chloride and potassium oxalate and then purified via recrystallization, a common technique in synthetic procedures. The exact chemical formula of the crystal will initially be unknown, though the general formula can be written as  $K_wFe_x(C_2O_4)_y(H_2O)_z$ . Titration and spectrophotometry will be used analyze the purified materials in order to determine the precise formula of the crystal.

### Part 1: Synthesis of Green Crystal

Aqueous solutions of iron (III) chloride hexahydrate and potassium oxalate monohydrate will be reacted at a temperature of around 90°C to produce the Green Crystal. The unbalanced chemical equation for the process may be written as follows:



Cooling the reaction mixture will result in the formation of green crystals. The crystals obtained in this step are considered as the "crude product" as they are not pure. The solid product will be separated from the aqueous supernatant by

decanting the liquid, and then the crude product will be purified in the next step of the process.

## **Part 2: Purification of the Green Crystal**

The purification of the Green Crystal is done through a process called **recrystallization**.

In recrystallization, the crude product is dissolved in a small amount of hot solvent, completely dissolving the entire sample. At this stage both the impurities and the desired product are dissolved. The choice of solvent is such that, the pure product is much less soluble in the cold solvent, causing them to solidify (recrystallize) as the solution cools, while the impurities should remain dissolved. The liquid, containing the impurities, can then be removed via filtration, leaving the pure solid product behind.

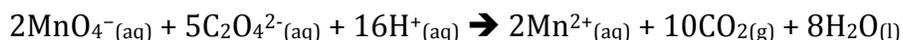
In this experiment, the most likely impurities combined with the pure product are the starting materials. According to the solubility rules, chloride salts (such as  $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$ ) and potassium salts (such as  $\text{K}_2\text{C}_2\text{O}_4 \cdot \text{H}_2\text{O}$ ) are highly soluble in water. The desired product, the Green Crystal ( $\text{K}_w\text{Fe}_x(\text{C}_2\text{O}_4)_y \cdot z\text{H}_2\text{O}$ ) is a sparingly soluble salt that only dissolves in water at high temperatures. Therefore, water is a good solvent choice for the recrystallization process.

The quality of the crystals is improved if the cooling is done slowly. Given this, it is not recommended that the cooling be done over ice initially. It is best to allow the solution to cool to room temperature slowly and then place the solution in an ice bath only if necessary.

The percentage purity of the recrystallized product can be greatly increased if the process is repeated more than once. While multiple recrystallizations increase the quality of the product, each iteration also decreases the quantity of product that can be isolated, resulting in an overall diminished percent yield. Given this concern only one or two recrystallizations are typically carried out in a typical synthetic scheme.

## **Part 3: Analysis of the Percent of Oxalate in the Green Crystal**

The percent of oxalate ( $\text{C}_2\text{O}_4^{2-}$ ) in the Green Crystal will be analyzed by titration (see experiments A7 and A9 for a review of titration). Oxalate reacts with permanganate in an acidic medium according to the following equation:



The above reaction is an example of a redox titration. The  $\text{MnO}_4^-$  is reduced to  $\text{Mn}^{2+}$  and the  $\text{C}_2\text{O}_4^{2-}$  is oxidized to  $\text{CO}_2$ . An introduction to oxidation-reduction (redox) reactions can be found in the Types of Reactions chapter of any General Chemistry textbook, while a more detailed treatment may be found in the later

electrochemistry chapter. Both of these sections will be discussed over the course of the General Chemistry series.

The primary focus for the present is the fact that according to the stoichiometry of the reaction, 2 moles of  $\text{MnO}_4^-$  will completely react with 5 moles of  $\text{C}_2\text{O}_4^{2-}$ .

During this titration, a known mass of the Green Crystal will be dissolved in warm acidic water. This solution will be colorless initially. A standardized permanganate solution (whose molarity is precisely known) will be slowly added from a buret into a solution in which some of the Green Crystal was dissolved. The purple color of the permanganate will disappear as it reacts with the oxalate in the reaction flask. Just past the equivalent point, when all the oxalate has completely reacted, addition of one more drop of the permanganate will leave an excess of permanganate in the solution and therefore a pale purple color in the reaction flask. This is the end point of the titration.

The volume of permanganate solution needed for the titration and the molarity of the permanganate will be used to determine the moles of permanganate used for the titration. From the stoichiometric relationship between the permanganate and the oxalate, the moles of the oxalate titrated can then be determined. This is then used to calculate the mass percent of oxalate in the sample.

#### **Part 4: Analysis of Percent of Iron in Green Crystal**

The analysis of the percentage of iron in the Green Crystal will be done using the method of spectrophotometry.

Spectrophotometry and Beer's law have been extensively discussed in previous experiments (see Experiments B4 and B6).

The iron in the Green Crystal has an oxidation state of +3. The analysis is greatly facilitated by converting all the Fe(III) to Fe(II). In order to reduce the iron from its +3 to +2 oxidation state, several steps are required. The sample is first placed in an ammonium acetate solution, which acts as a buffer to stabilize the pH of the solution. The reduction of Fe(III) to Fe(II) is done by adding hydroxylamine hydrochloride. Finally, the reagent ortho-phenanthroline is added to the Fe(II) to give the solution an orange color. Ortho-phenanthroline acts as a Lewis base here, donating electron density to the iron cation, forming the Lewis adduct  $[\text{Fe}(\text{phen})_3]^{2+}$ , called "ferroin." This orange colored complex is commonly used for the photometric determination of Fe(II) concentrations.

The orange colored ferroin complex has a wavelength of maximum absorbance ( $\lambda_{\text{max}}$ ) at 510 nm. Therefore, spectrophotometric analysis of the sample is carried out at 510 nm.

In the previous discussions of spectrophotometry and Beer's law, the calculation of the concentration of a sample whose absorbance is known uses the formula:

$$A = \epsilon \times C \times l$$

In the above formula: A is the absorbance of the sample, C is the concentration, l is the path length of the instrument and  $\epsilon$  is the molar extinction coefficient of the substance being analyzed.

The above formula can only be used if the molar extinction coefficient of the sample is known. In the absence of this information, additional steps must be employed to determine the concentration of the unknown sample. One of the simplest methods involves constructing a standard curve using known concentrations of the unknown. This standard curve is a plot of the absorbance of known concentrations of the analyte, and is often referred to as a "Beer's Law Plot."

A plot of absorbance (y-axis) vs. concentration (x-axis) should result in a curve whose regression equation corresponds to that of a straight line,  $y = mx + b$ . Upon measuring the absorbance of the unknown sample, the regression equation obtained from the standard curve will be used to calculate the concentration of the unknown (concentration  $x = (y - b)/m$ ).

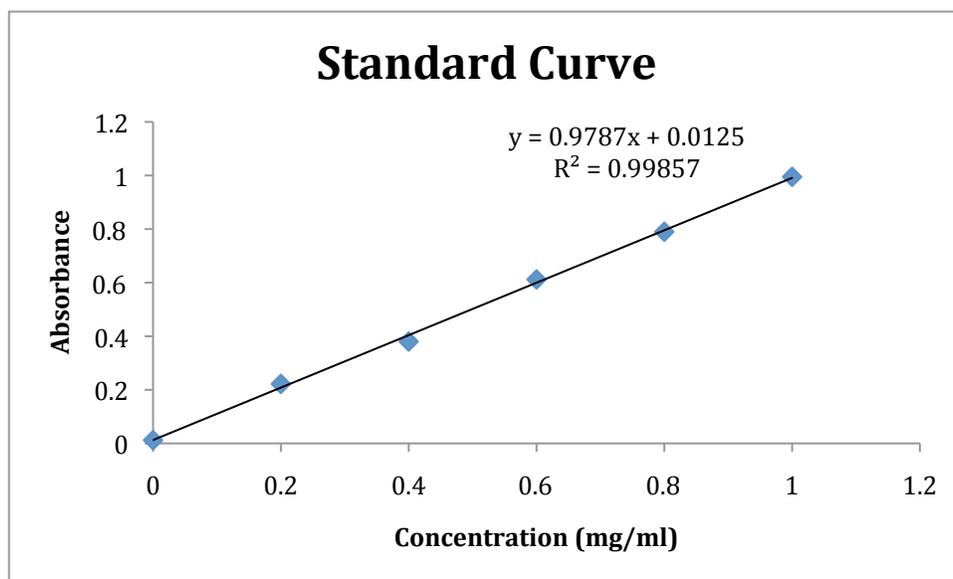
In effect, the slope value obtained in this graph should equal the product of the path length and the molar extinction coefficient,  $\epsilon \times l$ . If desired, the value of the molar extinction coefficient can be determined at this point as well by dividing the slope by the path length.

By way of example as to how this would be employed in an experiment, assume that a stock solution of the analyte with a concentration of 1 mg/ml has been provided.

The known solution can be diluted to obtain a set of solutions of known concentrations such as: 0.2, 0.4, 0.6, 0.8, and 1.0 mg/ml. The absorbance measurement of each of these solutions and the unknown solution may result in a data set such as is presented below:

<b>Solution (mg/ml)</b>	<b>Absorbance</b>
0.0	0.012
0.2	0.222
0.4	0.380
0.6	0.612
0.8	0.790
1.0	0.995
Unknown	0.585

In order to ascertain the concentration of the unknown, a standard curve must be constructed. As mentioned before, the standard curve is a plot of absorbance (y-axis) vs. concentration (x-axis). The graph below is the standard curve obtained from the above data. The equation of the best-fit line (see Experiment B2) is also shown in the graph.



Since the absorbance of the unknown sample is 0.585, its concentration is calculated using the equation of the line as follows:

$$y = 0.9787x + 0.0125$$

$$x = \frac{y - 0.0125}{0.9787}$$

$$x = \frac{0.585 - 0.0125}{0.9787} = 0.58 \approx 0.6 \frac{mg}{ml}$$

### **Part 5: Analysis of Percent Hydrate in Green Crystal**

From the formula of the Green Crystal,  $K_wFe_x(C_2O_4)_y \cdot zH_2O$ , it should be apparent that the substance is a hydrate. As was discussed in Experiment A3, hydrates are inorganic compounds with water molecules bound to them. In a hydrated compound a fixed number of water molecules, called the water of hydration, are chemically combined with the metal ion.

In this part of the analysis, the mass percent of water in the hydrated Green Crystal will be determined using methods originally described in Experiment A3. The

procedure involves measuring the mass of the substance before and after it is dehydrated by heating in order to determine the mass of water lost in the process.

### **Part 6: Analysis of Percent Potassium in Green Crystal**

The determination of mass percents of iron, oxalate, and water was discussed in Parts 3, 4, and 5, respectively. The mass percent of potassium is then calculated by difference. No separate analytical procedure is needed for this determination.

## Experimental Design

This experiment will be performed over a period of four laboratory days.

**Day 1:** Iron (III) chloride hexahydrate and potassium oxalate monohydrate will be combined in an aqueous medium at a temperature of 90°C and cooled to obtain the crude product of the Green Crystal. The crude product will then be recrystallized using water as the solvent and cooled overnight to obtain the pure form of the Green Crystal.

**Day 2:** The mass percent of oxalate in the purified Green Crystal will be determined by titration with standardized potassium permanganate.

**Day 3:** A Beer's Law Plot will be constructed and the mass percent of iron in the purified Green Crystal will be determined by spectrophotometric analysis.

**Day 4:** The mass percent of water in the purified Green Crystal will be determined by measuring the mass of the hydrated and dehydrated salt.

## Reagents and Supplies

Solid Iron (III) chloride hexahydrate, solid potassium oxalate monohydrate, 0.04 M Standardized  $\text{KMnO}_4$ , 3 M  $\text{H}_2\text{SO}_4$ , Standard iron solution (0.05 mg/ml), 10% hydroxylamine hydrochloride, 3% o-phenanthroline, 1 M ammonium acetate solution

Filter paper, filter flask, Hirsch funnel, 25-mL buret, hot plate, seven 10-mL volumetric flasks, two 25-mL volumetric flasks, Spectrophotometer, cuvettes, micro crucible

(See posted Material Safety Data Sheets)

## Procedure

### PART 1: SYNTHESIS OF GREEN CRYSTAL

1. Obtain a hot plate from the stock room.
2. Measure 3-4 grams of iron (III) chloride hexahydrate and transfer it to a 250-mL beaker. Record the exact mass.
3. Add 15 mL of deionized water to the iron (III) chloride hexahydrate to dissolve. Warm slightly if necessary.
4. Measure 8-9 grams of potassium oxalate monohydrate and transfer it to a 250-mL beaker. Record the exact mass.
5. Add 30 mL of deionized water to the potassium oxalate monohydrate to dissolve. Heat the solution to boiling.
6. Transfer the iron (III) chloride hexahydrate solution to the beaker containing the boiling potassium oxalate monohydrate.
7. Heat the mixture to about 90°C.
8. When the mixture reaches 90°C, from heat using tongs. Place the beaker on a red tile and allow it to cool slowly for five minutes.
9. Prepare a tray with ice.
10. Place the beaker containing the mixture of reagents on ice and allow it to cool for 30-40 minutes.
11. Once crystallization is complete, slowly decant the supernatant and discard the liquid into an appropriate waste container provided by the instructor.
12. The solid crystals are the crude product. Take these forward to the next step.

PART 2: PURIFICATION OF THE GREEN CRYSTAL

1. Heat approximately 25 mL of deionized water to near boiling.
2. Add a couple of mL of hot water from Step 1 to the crude product from Part 1 and heat this mixture along with the deionized water. Continue adding small increments of hot water to the mixture containing the crude product until the crude product is completely dissolved. Make sure not to add too much water. [NOTE: In order for the recrystallization to be successful, the crude product must be dissolved in a minimum amount of hot solvent.]
3. Cover the beaker with a watch glass.
4. Place the beaker containing the dissolved crude product inside the locker. Purified crystals will form as the solution cools.
5. This is a good stopping point for Day 1 activities.

**Day 2**

6. Cool a squirt bottle containing deionized water in an ice bath.
7. Obtain the beaker containing the pure Green Crystals from the locker.
8. Set up a vacuum filtration using a Hirsch funnel and a filter flask.
9. Transfer the crystals from the beaker to the filter paper on the Hirsch funnel. Scrape any crystals stuck to the bottom of the beaker.
10. Rinse the beaker with two small aliquots of ice-cold water and add this to the crystals.
11. Remove the filter paper containing the crystals from the Hirsch funnel and place on a watch glass.
12. Allow the crystals to air dry.
13. While the crystals are drying start working on Part 3 and return to attend to the crystals when they are completely dry.
14. Obtain the mass of the dry crystals. Be sure to weigh all the crystals synthesized.
15. Transfer the crystals into a labeled beaker and place them in the locker until the end of the experiment.

PART 3: ANALYSIS OF THE PERCENT OXALATE IN THE GREEN CRYSTAL

1. Obtain about 20 mL of standardized  $\text{KMnO}_4$  solution. Record the exact molarity of the  $\text{KMnO}_4$ .
2. Obtain a 25-mL buret and a hot plate from the stockroom.
3. Set up a buret stand and clamp the buret to the stand. Turn on the hot plate.
4. Rinse the buret twice with deionized water and condition the buret with  $\text{KMnO}_4$ .
5. Fill the buret with the  $\text{KMnO}_4$  solution and record the initial buret reading.
6. Measure 0.1 to 0.12 grams of pure Green Crystals. Record the exact mass.
7. Transfer the crystals to a 125-mL Erlenmeyer flask.
8. Add about 12.5 mL of deionized water and 2.5 mL of 3 M sulfuric acid to the Erlenmeyer flask. Add a boiling chip into the flask.
9. Gently heat the Erlenmeyer flask on the hot plate until the crystals dissolve completely.
10. Place the Erlenmeyer flask containing the hot solution below the buret containing the  $\text{KMnO}_4$  solution and begin titrating. Constantly swirl the flask to thoroughly mix the contents. Stop the titration at the appearance of a permanent pale purple color.
11. Record the final buret reading.
12. Repeat steps 6 to 12 an additional two to three times, as needed. You should have 2 trials in which the mass percent agrees within 10%.
13. Discard all the waste in an appropriate waste container provided by the instructor.

PART 4: ANALYSIS OF PERCENT IRON IN GREEN CRYSTAL

NOTE: This part of the experiment will be performed in groups of two. The instructor will obtain the data for the standard curve that will be used by the entire class. Both members of the group must prepare their own sample for analysis.

1. Each group should obtain two 10-mL volumetric flasks and two 25-mL volumetric flasks from the stockroom.
2. Obtain two cuvettes and a spectrophotometer. Power up the spectrophotometer and allow the instrument to warm up for 10 minutes. The instructor will demonstrate the proper use of the spectrophotometer.

*Preparation of Standards (Steps 3-9 will be done by the instructor)*

3. Label five 10-mL volumetric flasks as 1, 2, 3, 4, and 5 and add the following reagents into the flask using electronic pipets for dispensing the reagents.

Flask	Iron-standard, $\mu\text{L}$	Ammonium acetate, $\mu\text{L}$	Hydroxylamine hydrochloride, $\mu\text{L}$	O-phenanthroline, $\mu\text{L}$
1	100	100	100	100
2	200	100	100	100
3	300	100	100	100
4	400	100	100	100
5	500	100	100	100

4. Fill the volumetric flasks with deionized water until the calibration mark. If too much water has been added, discard this solution in the waste container and redo the preparation.
5. Cap the volumetric flask, ensure that the cap is secure, and thoroughly mix the contents of the flask by inverting.
6. Prepare five "Blank" solutions corresponding to each standard solution by substituting the Iron-standard with an equivalent amount of deionized water.
7. Measure the absorbance of the samples at 510 nm.
8. Add the blank solution to one of the cuvettes. Place the cuvette containing the blank inside the sample compartment, align the guide mark on the cuvette with the guide mark at the front of the sample compartment, close the lid and adjust the absorbance reading to zero.

9. Add the standard from volumetric flask 1 to a cuvette. Now place the cuvette containing the standard inside the sample compartment, align the guide mark on the cuvette with the guide mark at the front of the sample compartment, close the lid and record the absorbance.

*Preparation of Samples (individually)*

10. Measure 0.01 grams of the Green Crystal. Record the exact mass of the Green Crystal. Transfer the crystals to a small beaker.
11. Add 10 ml of deionized water to the beaker.
12. Add 0.5 ml of 3 M sulfuric acid, using a microburet.
13. Heat the beaker till all the salt dissolves.
14. Transfer the contents into a 25-mL volumetric flask. Rinse the beaker with deionized water and transfer the rinse solution into the volumetric flask as well.
15. Fill the volumetric flask with deionized water until the calibration mark. If too much water has been added, prepare the solution in this flask again.
16. Cap the volumetric flask, ensure that the cap is secure, and thoroughly mix the contents of the flask by inverting.
17. Carefully transfer 0.50 mL of the above solution into a 10-mL volumetric flask labeled as "sample". Use a microburet for the measurement.
18. To the 10-mL volumetric flask containing the sample also add (use electronic pipets):
  - a. 100  $\mu$ L ammonium acetate
  - b. 100  $\mu$ L hydroxylamine hydrochloride
  - c. 100  $\mu$ L O-phenanthroline
19. Fill the volumetric flask with deionized water until the calibration mark. If too much water has been added, prepare the solution in this flask again.
20. Cap the volumetric flask, ensure that the cap is secure, and thoroughly mix the contents of the flask by inverting. This will be the "Sample".

*Measurement of Absorbance*

21. Set the wavelength of the spectrophotometer to 510 nm

22. In a large test tube prepare a "Blank" solution by adding the following, followed by thorough mixing:
- 100  $\mu\text{L}$  ammonium acetate
  - 100  $\mu\text{L}$  hydroxylamine hydrochloride
  - 100  $\mu\text{L}$  O-phenanthroline
  - 10 mL of deionized water
23. Measure the absorbance of the "Sample" (from Step 20) according to methods described in Steps 7-9. Use the "Sample" instead of the "Standards" for the measurement.

PART 5: ANALYSIS OF PERCENT HYDRATE IN GREEN CRYSTAL

1. Obtain a micro crucible from the stockroom.
2. Heat an empty micro crucible and bring it to a constant mass ( $\pm 0.0010$  g). In order to do this, record the mass of the empty crucible. Then heat the crucible for 3-4 minutes. Once the crucible is cool, record its mass again. Repeat the heating-cooling-weighing process until the mass recorded is within 1 mg of the previous measurement.
3. Add some Green Crystal to the micro-crucible (scoop a small amount with a spatula, approximately 0.1000 grams). Measure the mass of the micro-crucible with the solid.
4. Heat the micro-crucible over a Bunsen burner (use a yellow flame). Continue the heating process for about five minutes. Make sure that the salt does not turn black in color (which is an indication of degradation of the product).
5. Cool the micro-crucible to room temperature and measure the mass of the micro-crucible with the anhydrous compound.
6. In order to ensure that all the water of hydration has been evaporated from the solid, the micro-crucible containing the anhydrous solid should be heated over the flame once again.
7. Once again, cool the micro-crucible and measure the mass of the micro-crucible containing the anhydrous compound.
8. The heating-cooling-weighing of the micro-crucible with the anhydrous compound must be repeated until the mass of the micro-crucible containing the anhydrous compound is within 1 mg of a previous measurement.
9. Discard the contents of the micro-crucible in the appropriate waste disposal container.

## Data Table

### PART 1: SYNTHESIS OF GREEN CRYSTAL

Mass of Iron (III) chloride hexahydrate, grams	
Mass of potassium oxalate monohydrate, grams	

### PART 2: PURIFICATION OF THE GREEN CRYSTAL

Mass of pure Green Crystal synthesized, grams	
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### PART 3: ANALYSIS OF THE PERCENT OXALATE IN THE GREEN CRYSTAL

Molarity of standardized $\text{KMnO}_4$			
	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>
Initial buret reading of $\text{KMnO}_4$ (mL)			
Final buret reading of $\text{KMnO}_4$ (mL)			
Mass of pure Green Crystals (grams)			

PART 4: ANALYSIS OF PERCENT IRON IN GREEN CRYSTAL

Mass of Green Crystal (grams)		
Concentration of stock iron solution		
	<b>Absorbance</b>	<b>Absorbance</b>
<b>Standard 1</b>		
<b>Standard 2</b>		
<b>Standard 3</b>		
<b>Standard 4</b>		
<b>Standard 5</b>		
<b>Sample</b>		

PART 5: ANALYSIS OF PERCENT HYDRATE IN GREEN CRYSTAL

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>
Constant mass of empty micro-crucible (grams)			
Mass of micro-crucible + Green Crystal (grams)			
Mass of micro-crucible + anhydrous Green Crystal (following first heating) (grams)			
Mass of micro-crucible + Green Crystal (following second heating) (grams)			
Mass of micro-crucible + Green Crystal (following third heating) (grams)			
Mass of micro-crucible + Green Crystal (following fourth heating) (grams)			

## Data Analysis

### PART 3: ANALYSIS OF THE PERCENT OXALATE IN THE GREEN CRYSTAL

Molarity of standardized $\text{KMnO}_4$			
	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>
Initial buret reading of $\text{KMnO}_4$ (mL)			
Final buret reading of $\text{KMnO}_4$ (mL)			
Volume of $\text{KMnO}_4$ (ml)			
Volume of $\text{KMnO}_4$ (L)			
Moles of $\text{KMnO}_4$ (Molarity $\times$ Volume)			
Moles of oxalate (use stoichiometry)			
Molar Mass of oxalate (grams/mole)			
Mass of oxalate (grams)			
Mass of pure Green Crystals (grams)			
Mass percent of oxalate			
Average mass percent of oxalate			

PART 4: ANALYSIS OF PERCENT IRON IN GREEN CRYSTAL

1. Calculate the concentration of iron in each of the Standard samples 1 through 5.

Concentration of stock iron solution =

NOTE: Each sample was diluted to a final volume of 10.0 mL. For instance, Standard 1 was prepared by diluting 100  $\mu\text{L}$  or 0.10 mL of stock solution to a final volume of 10.0 mL

	<b>Concentration of Iron</b>	<b>Average Absorbance</b>
Standard 1		
Standard 2		
Standard 3		
Standard 4		
Standard 5		
Sample	Unknown	

2. Plot the Standard Curve (Plot of absorbance (y-axis) vs. concentration (x-axis)) and obtain the equation of the best-fit line.

Equation of best-fit line:



PART 5: ANALYSIS OF PERCENT HYDRATE IN GREEN CRYSTAL

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>
Mass of empty micro-crucible (grams)			
Mass of micro-crucible + Green Crystal (grams)			
Mass of Green Crystal (grams)			
Constant mass of micro-crucible + anhydrous Green Crystal (following final heating) (grams)			
Mass of anhydrous Green Crystal (grams)			
Percentage of water in Green Crystal (grams)			
<b>Average percentage of water in Green Crystal</b>			

PART 6: ANALYSIS OF PERCENT POTASSIUM IN GREEN CRYSTAL

Percent oxalate	
Percent iron	
Percent hydrate	
Therefore percent potassium	

PART 7: DETERMINE THE EMPIRICAL FORMULA OF THE GREEN CRYSTAL

Assume there are 100 grams of the Green Crystal:

	<b>Grams</b>	<b>Molar Mass (grams/mol)</b>	<b>Moles</b>	<b>Relative Moles</b>
<b>Oxalate (C<sub>2</sub>O<sub>4</sub><sup>2-</sup>)</b>				
<b>Iron</b>				
<b>Hydrate (H<sub>2</sub>O)</b>				
<b>Potassium</b>				

Empirical Formula of Green Crystal:
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