1. What is the purpose of this experiment?

2. Write the generic chemical formula for the coordination compound (complex salt) you made last week. (Do not write the chemical reaction. You will not receive any credit for writing the chemical reaction.)

3. When 0.1009 g of sodium cyanide is dissolved in solution, it takes 19.22 mL of a potassium permanganate solution to reach the end point of the titration. Using the balanced chemical equation provided, determine the molarity of the potassium permanganate. You will be doing a similar reaction in the laboratory this week. (Remember: the mole ratio of \( \text{MnO}_4^- \) to \( \text{KMnO}_4 \) is 1:1 and the mole ratio of \( \text{CN}^- \) to \( \text{NaCN} \) is 1:1)

\[
\text{H}_2\text{O} \ (\ell) \ + \ 2 \text{MnO}_4^- \ (aq) \ + \ 3 \text{CN}^- \ (aq) \ \rightarrow \ 2 \text{MnO}_2 \ (s) \ + \ 3 \text{CNO}^- \ (aq) \ + \ 2 \text{OH}^- \ (aq)
\]
Coordination Compounds with Copper (II)
Prelab (Week 3)

Name___________________________________ ____________________
Total________/10

SHOW ALL WORK       NO WORK = NO CREDIT

1. You weigh out 0.1183 g of a complex salt to analyze for the percentage of cyanide ion in your complex salt. After dissolving the complex salt in solution, you determine that it takes 10.02 mL of a 0.08035 M potassium permanganate solution to react with all of the cyanide ion in your complex salt. Using the balanced chemical equation provided, determine the mass percentage of the cyanide ion in your complex salt. (Remember: the mole ratio of MnO$_4^-$ to KMnO$_4$ is 1:1)

\[
\text{H}_2\text{O} (\ell) \quad + \quad 2 \text{MnO}_4^- (\text{aq}) \quad + \quad 3 \text{CN}^- (\text{aq}) \rightarrow \quad 2 \text{MnO}_2 (\text{s}) \quad + \quad 3 \text{CNO}^- (\text{aq}) \quad + \quad 2 \text{OH}^- (\text{aq})
\]

2. You want to determine the amount of iron(III) in a complex salt. You make a 100.0 mL solution of the complex salt by dissolving 0.0670 g of complex salt with water to the mark on a 100.0 mL volumetric flask. You get an absorbance reading for the iron from the spectrometer of 0.334. The molar absorptivity is determined to be 17605 L/mole∙cm and the pathlength is 1.00 cm. Determine the molarity of iron(III) in solution. Determine the mass percentage of iron(III) in the complex salt.
Coordination Compounds with Copper (II)

You will synthesize a complex salt containing copper and then analyze it using redox volumetric methods and spectrometry to determine the percent composition and ultimately, its empirical formula. You will learn the important technique of titration. There will be two parts to this laboratory experiment: Part I: the synthesis of the complex salt and Part II: the analysis of the complex salt. The laboratory experiment will be performed over a three week period.

Special mention goes to John Curtin and Nevin Gunduz for developing this laboratory experiment.

Introduction

Ions of the transition elements exist in aqueous solution as complex ions. A complex ion is a metal ion with a molecule called a ligand attached to it through coordinate covalent bonds. A complex salt is an ionic compound but it differs in the fact that there are these covalent bonds attaching the metal to the ligand. The iron(II) ion, for example, exists in water as Fe(H₂O)₆²⁺. The water molecules in this ion are arranged about the iron atom with their oxygen atoms bonded to the metal by donating electron pairs to it. Replacing the H₂O molecules with six CN⁻ ions gives the Fe(CN)₆⁴⁻ ion. A coordination compound is a compound consisting either of complex ions and other ions of opposite charge or a neutral species. An example of a coordination compound is the compound K₄[Fe(CN)₆] which contains the complex ion Fe(CN)₆⁴⁻ and four K⁺ ions. Some of the transition elements have biological activity and their role in human nutrition depends in most cases on the formation of complexes and coordination compounds.

In the first part of this experiment, you will make a complex salt containing copper in which water molecules and oxalate ions (C₂O₄²⁻) act as ligands. This complex salt will be blue in colour. The formula for this complex will have the formula KₓCuᵧ(C₂O₄)₂•2H₂O where K has a charge of +1, Cu has a charge of +2 and C₂O₄ has a charge of -2. After this salt has been prepared you will determine its percent composition in copper and oxalate ion and from this information, determine the values for x, y and z in the formula written above.

In the synthesis of the complex salt, you will dissolve a copper(II) compound and form the coordination compound. When your coordination compound is crystallized, you can begin the analysis of this compound for the percentage of oxalate and copper(II). The overall reaction for this synthesis is:

\[ y \text{CuSO}_4 \cdot 5\text{H}_2\text{O} + \text{excess reagent} \rightarrow K_x\text{Cu}_y(\text{C}_2\text{O}_4)_z \cdot 2\text{H}_2\text{O} + \text{other products} \]

It is assumed that CuSO₄ • 5H₂O is the limiting reagent and when the number of moles of copper (y) are determined from the percent composition of the coordination compound above, the equation can be balanced and the theoretical yield of this compound can be determined.

In order to determine the amount of oxalate in your compound, you must know the concentration of the potassium permanganate (KMnO₄) solution you are going to use for Part IIIB. You must “standardize” or determine the concentration of a KMnO₄ solution in Part IIA. You will perform the standardization of a KMnO₄ solution by titrating a permanganate solution into a solution containing a known mass of sodium oxalate (Na₂C₂O₄). The titration reaction is:

\[ 5 \text{C}_2\text{O}_4^{2-} (\text{aq}) + 2 \text{MnO}_4^- (\text{aq}) + 16 \text{H}^+ (\text{aq}) \rightarrow 10 \text{CO}_2 (\text{g}) + 2 \text{Mn}^{2+} (\text{aq}) + 8 \text{H}_2\text{O} (\text{l}) \]

A titration is defined as a procedure for the quantitative analysis of a substance by means of an essentially complete reaction in solution with a reagent of known concentration. The reagent of known concentration in our case is Na₂C₂O₄. A buret will be used to perform the titration. The buret holds one of the reactants called the titrant and adds this into a reaction vessel which contains the other reactant. The titrant in this experiment
will be the KMnO₄ solution that you make. The Na₂C₂O₄ solution will be in the reaction vessel. When the reaction is complete, the solution will be a light pink colour. At some point in your titration, you will have added the same number of moles of KMnO₄ as there are moles of Na₂C₂O₄ in the reaction vessel. In other words, there will be equal number of moles of KMnO₄ and Na₂C₂O₄ in the solution. This point in the reaction is called the equivalence point or endpoint. At this point, the solution in the reaction vessel will change colour when the slightest amount of titrant is added. It is at this point you want to stop and record this final volume.

At the completion of the reaction, the moles of KMnO₄ which you added from the buret can be determined from the grams of Na₂C₂O₄ you weighed out. You can convert the grams of Na₂C₂O₄ to moles and then determine the moles of KMnO₄ from the stoichiometry of the reaction. It is important to remember that the moles of Na₂C₂O₄ are the same as the moles of C₂O₄²⁻ and the moles of KMnO₄ are the same as the moles of MnO₄⁻.

The molarity of KMnO₄ can then be calculated by using the moles of KMnO₄ which you calculate and the volume of KMnO₄ that was added from the buret.

In Part IIC of this experiment, you will be using spectroscopy to determine the percentage of copper(II) in your complex salt. You will use the theory of Beer’s Law to determine a concentration of copper(II) in a solution and then from that information, determine the mass percentage of copper in your sample.

There are many ways to determine concentrations of a substance in solution. There are other properties of a solution that change with concentration such as colour. Beer’s law relates colour intensity and concentration. It states that concentration and absorbance are linearly related. This means that the higher the concentration, the higher the absorbance and vice versa. Using colour can be much faster than using titration, especially when you have many samples containing different concentrations of the same substance. When coloured solutions are irradiated with white light, they will selectively absorb light of some wavelengths, but not of others. The wavelength at which absorbance is highest is the wavelength to which the solution is most sensitive to concentration changes. This wavelength is called the maximum wavelength or λmax. In this laboratory experiment, you will be taking the copper(II) in your complex salt and complexing it with ammonia to produce the copper(II)-ammonia complex which is dark blue in color. The chemical reaction for this complex is:

\[ \text{Cu}^{2+}(aq) + 4 \text{NH}_3(aq) \rightleftharpoons \text{Cu(NH}_3)_4^{2+}(aq) \]

In spectrophotometric studies, Beer’s Law can be stated as \( A = \varepsilon b c \) where \( \varepsilon \) is the molar absorptivity, \( b \) is the path length in cm, \( c \) is the concentration in moles/L and \( A \) is absorbance. At 619 nm, only the copper(II)-ammonia complex absorbs light and will be the wavelength you will use on your spectrometer. This means Beer’s Law may now be written as

\( A = \varepsilon b [\text{Cu(NH}_3)_4^{2+}] \) where \([\text{Cu(NH}_3)_4^{2+}]\) is the same as \([\text{Cu}^{2+}]\) which can then be rearranged to form

\[ [\text{Cu}^{2+}] = \frac{A}{\varepsilon b} \]

The molar absorptivity for the copper(II)-ammonia complex is 56.78 L/mole • cm. You will then be able to calculate the concentration for copper(II) from the absorbance value you will obtain in Part IIC.
BE SURE THAT YOU RECORD ALL OF YOUR DATA PROPERLY FOR ALL PARTS OF THIS LABORATORY EXPERIMENT. YOU WILL NOT BE RECEIVING ANY RESULTS SHEETS AS A GUIDE ON HOW TO RECORD YOUR DATA.

Part I
Synthesis of the Complex Salt containing Copper

Procedure

1. In a 250.0 mL beaker, weigh out approximately 10.00 g of potassium oxalate monohydrate, K₂C₂O₄ • H₂O on the top loader balance and add 20.00 mL of distilled water.

2. Heat this solution, while stirring, until all of the potassium oxalate has dissolved. Do not boil the solution. Stop heating the solution when the solution becomes clear.

3. While the potassium oxalate monohydrate solution is heating, you can make your solution of copper(II) sulfate pentahydrate (CuSO₄ • 5H₂O). In a 250.0 mL beaker, weigh out approximately 1.00 g of copper(II) sulfate pentahydrate. Be sure to use the top loader balance and record the mass to the nearest or 0.01g.

4. Dissolve the copper(II) sulphate pentahydrate in 15.00 mL of water.

5. Pour the potassium oxalate monohydrate solution into the copper(II) sulphate pentahydrate solution.

6. Allow to cool to room temperature, you should see crystals forming. Let sit for another 15 minutes after the solution has cooled to room temperature.

7. Check your suction filtration apparatus as shown by your instructor.

8. Use a rubber policeman to scrape the crystals and solution into the funnel. If you have any crystals that remain in the beaker, wash them into the funnel with some 1:1 alcohol (ethanol)-water solution.

9. Rinse the crystals while they are in the funnel with 20 mL of a 1:1 alcohol(ethanol)-water solution. Pour this solution slowly so all of the crystals are washed. Pour 10 mL of acetone over the crystals making sure that it is poured over all the crystals. Allow to dry while suction is running for another 5-10 minutes. Weigh a clean, dry 250 mL beaker and record its mass. Use the rubber policeman to scrape the crystals off the filter paper and filter funnel into the clean dry beaker and place in your drawer. Do not cover the beaker. You will weigh the beaker with your crystals next week.
Part II
Analyzing the Complex Salt

NOTE: Place the beaker with your dry crystals on the top loader balance and record the mass of the crystals you produced. **Record this mass since this is the experimental yield that you will need later to calculate the percent yield for the crystals.**

There are three sections to part II of this experiment. Part A involves standardizing your potassium permanganate solution which you will use in the next two parts of the experiment. Part B involves determining the amount of oxalate in the coordination compound and Part C involves analyzing the amount of copper in your coordination compound.

**Part II A: Standardizing a Potassium Permanganate Solution**

In order to determine the amount of oxalate in your compound, you must know the concentration of the potassium permanganate (KMnO₄) you are going to use in part B.

You will do this by titrating your permanganate solution into a solution containing a known mass of sodium oxalate (Na₂C₂O₄). The titration reaction is:

\[
5 \text{C}_2\text{O}_4^{2-} (\text{aq}) + 2 \text{MnO}_4^- (\text{aq}) + 16 \text{H}^+ (\text{aq}) \rightarrow 10 \text{CO}_2 (\text{g}) + 2 \text{Mn}^{2+} (\text{aq}) + 8 \text{H}_2\text{O} (\text{l})
\]

Data that you should have written down for EACH titration for Part II A are:

- Titration reaction
- Mass of Na₂C₂O₄ in sample
- Final Buret Reading
- Initial Buret Reading
- Volume of KMnO₄ used
- Moles of Na₂C₂O₄
- Moles of C₂O₄²⁻
- Moles of MnO₄⁻ used (look at titration reaction above)
- Moles of KMnO₄ used
- Molarity of KMnO₄
- Average Molarity of KMnO₄ (for all titrations that are good)

You should have sample calculations of just ONE titration shown in your laboratory notebook.

**Procedure**

**It is very important that all of your apparatus be clean and that distilled water is used in all solutions.**

For your titrations (in parts A and B), always have a wash bottle with distilled water in it ready to use.

1. Collect 50.0 mL of the potassium permanganate solution that you are going to standardize.

2. Using the **analytical** balance, weigh approximately 0.1200 g of sodium oxalate and place in a 250 mL Erlenmeyer flask. Add 10 mL of 6 M H₂SO₄ to the flask and this will help the sample to dissolve. Add 65 mL of water to the Erlenmeyer flask.
3. You must make sure that your buret is free of contamination since this can seriously affect your data. First clean and rinse the buret properly. When you are ready to begin the titration, rinse the buret with three 5 mL sample of the titrant (in this case, the KMnO₄ solution). Let this solution drain out the bottom by opening the stopcock. Make sure the whole interior of the buret has been sufficiently rinsed with the titrant solution. Discard the titrant used to rinse the buret and close the stopcock and put your fresh titrant in the buret. Fill the buret as close as you can to the top of the buret. Now open the stopcock fully and allow some of the titrant to be released. This eliminates any bubbles at the tip which can cause error in your data. Try to start your titration with the titrant level near the 0.00 mark. It does not have to be exact so do not waste a lot of your time being exact. Try to get at the 0.00 mark or just below it. Always record this initial volume. The technique for reading the values from a buret are the same as those for the graduated cylinder. The buret is marked by 1 mL lines and the subdivisions represent 0.1 mL. You can therefore, estimate the location of the meniscus to 0.01 mL.

4. Heat the solution of sodium oxalate in the Erlenmeyer flask to 80-90 °C. When you remove the thermometer to perform the titration, be sure to rinse the thermometer into the flask since you do not want to lose any of the sodium oxalate.

5. Record the initial reading on the buret, to the nearest 0.01 mL and begin to add the KMnO₄ solution to the flask but do not add too rapidly and be sure to swirl the solution.

If you add the KMnO₄ solution too rapidly, or do not swirl well, you may find you have some brown colouration in your solution. This is due to the formation of manganese(IV) oxide (MnO₂). If you have not added any more KMnO₄ than needed to reach the endpoint, the excess oxalate should reduce the MnO₂ momentarily. However, if you fail to swirl the sample and overshoot the endpoint while MnO₂ is formed, the titration is ruined and must be performed again.

You will momentarily see a spot of pink color when you add the titrant, but it will disappear as you swirl. When you notice that the pink color is more persistent before being swirled away, you know you are nearing your endpoint. Slow down the addition of titrant at this time. You may want to keep a preliminary final volume reading recorded just in case you overshoot with one more drop. You have reached the endpoint when one drop of titrant causes the whole solution to turn from colorless to a permanent very faint pink color which will not disappear when you swirl. Record this final volume from the buret. Leave any leftover solution in your buret since you will have to do this another two times. Just refill the buret back to the 0.00 mark or close to it. It is not necessary to have the same initial volume for each titration. Just realize that your endpoint will occur at a slightly different value if your initial volume is not the same as the first titration.

6. You should start to notice that as you are nearing the endpoint of your titration that the decolouration of the KMnO₄ takes longer and longer. At this time you should add the KMnO₄ more slowly, preferably drop by drop. When you have reached the endpoint, there will be a faint colour that persists in the solution. If you are afraid that you will overshoot the endpoint with one more drop, then record the volume before you add that one drop and that way if you do overshoot, at least you have the final volume recorded and you will not have to perform an additional trial of the titration.

7. Repeat this procedure two more times.
Part IIB: Determination of the amount of Oxalate in the Complex Salt

The determination of oxalate in the complex salt involves oxidation-reduction or redox titrations. The KMnO₄ will oxidize all of the oxalate to CO₂ which escapes. You will use the same chemical equation from part A to calculate the amount of oxalate in the complex salt.

Data that you should have written down for EACH titration for Part IIB are:

- Titration reaction
- Mass of Complex Salt
- Final Buret Reading
- Initial Buret Reading
- Volume of KMnO₄ used
- Average Molarity of KMnO₄
- Moles of KMnO₄
- Moles of MnO₂⁻ used
- Moles of C₂O₄²⁻ in sample (use same chemical equation as Part IIA)
- Mass of C₂O₄²⁻ in sample
- Mass % of C₂O₄²⁻ (for all titrations that are good)

You should have sample calculations of just ONE trial shown in your laboratory notebook.

Procedure

1. Using the analytical balance, weigh out 0.1000-0.1500g of your complex salt and place in a 250 mL Erlenmeyer flask.

2. Dissolve the sample in 10 mL of 6 M H₂SO₄. You may need to warm this solution to dissolve your sample.

3. Add 65 mL of distilled water to the Erlenmeyer flask. Heat this solution to 80-90 °C and titrate as you did in part A. The endpoint will now have a pinkish-purple colour instead of a pink colour and this is because of the copper(II) ions which when in solution give a blue colour.

4. Repeat this procedure two more times.

Part IIC: Determination of the amount of Copper in the Complex Salt

Data that you should have written down for Part IIC are:

- Mass of complex salt
- Absorbance reading for copper-ammonia complex
- Concentration of the Cu(II) ion in the solution
- Moles of Cu(II) in sample
- Mass of Cu(II) in sample
- Mass % of Cu(II)

Calculations are to be shown in your laboratory notebook.
**Procedure**

1. Weigh out 0.1000 – 0.1300 g of your complex salt and place the salt in a 250.0 mL beaker.

2. Add 50.0 mL of 4 M NH₃ to the beaker, swirl until the crystals dissolve and transfer to a 100.0 mL volumetric flask. Your solution should have a blue colour.

3. Rinse your beaker with distilled water and pour into the volumetric flask.

4. Dilute to the 100.0 mL mark with distilled water.

5. Measure 5 mL of 4 M NH₃ and place in a small beaker. Then add 5 mL of distilled water to the same beaker and mix. This will make a 2 M NH₃ solution. You will use this solution as your blank solution.

6. When you are ready to read your absorbance value, open up the net book and click on logger pro. Click on the configure spectrometer icon on the tool bar and select absorbance vs. concentration.

   You then need to select your wavelength which will be around 619 nm (you may be off 1-0.5 nm, that’s okay). Below is what it should look like after you are done selecting the collection mode and wavelength.

   ![Collection Mode](image)

   Click okay and now you are ready to calibrate the spectrometer with the blank solution.

7. On the menu bar, select Experiment, then Calibrate, and then finally Spectrometer. Wait for the lamp to warm up.

8. Place your blank in the spectrometer and click Finish Calibration. When the calibration is finished, click OK. You are now ready to analyze the solution containing your complex salt.

9. Place the cuvette that contains the solution of the complex salt in the spectrometer and read the absorbance value in the bottom left part of the screen of the net book. Record this absorbance value.
CALCULATION EXAMPLE FOR PART IIC

The following is an example of how to do calculations for Part IIC. This is an example; you must use your own data when doing these calculations for your part of the experiment.

Mass of complex salt – 0.1631g

Absorbance value – 0.250

\[ \varepsilon = 56.78 \text{ L/mole} \cdot \text{cm} \quad b = 1 \text{ cm} \]

\[ A = \varepsilon bc \]

\[ c = \frac{A}{\varepsilon b} = \frac{0.250}{(56.78 \text{ L/mole} \cdot \text{cm})(1 \text{ cm})} = 0.00440 \text{ M} \]

This concentration value (0.00440 M) represents the concentration of Cu\(^{2+}\) in the volumetric flask. You can now calculate the number of moles of Cu\(^{2+}\) in the solution.

\[
0.1000 \text{ L} \times \frac{0.00440 \text{ moles of Cu(II)}}{1 \text{ L}} \times \frac{63.546 \text{ g Cu(II)}}{1 \text{ mole Cu(II)}} = 0.0280 \text{ g of Cu}^{2+}
\]

This represents the mass of Cu\(^{2+}\) in the 0.1631 gram sample of complex salt.

\[
\%\text{Cu}^{2+} = \frac{0.0280 \text{ g Cu(II)}}{0.1631 \text{ g of complex salt}} \times 100 = 17.2\% \quad \text{Cu}^{2+} \text{ in the complex salt}
\]

Calculation for Empirical Formula and Percent Yield

Once you have determined the percentages of oxalate and copper in your complex salt, you can determine the empirical formula of your salt and then the theoretical and percent yield of your coordination compound. The following is a sample calculation for empirical formula and theoretical and percent yield. THIS IS ONLY AN EXAMPLE TO HELP YOU WITH THE CALCULATIONS, YOU MUST USE YOUR OWN DATA.

Suppose you determine the following mass percentages: Mass % Oxalate = 55.10%  
Mass % Cu = 13.2%

Assume 100 g of complex salt.

Convert each mass to moles:

\[
\text{moles of oxalate} = \frac{55.10 \text{ g of oxalate}}{88.02 \text{ g of oxalate}} \times \frac{1 \text{ mole of oxalate}}{88.02 \text{ g of oxalate}} = 0.6260 \text{ moles of oxalate}
\]

\[
\text{moles of copper} = \frac{13.2 \text{ g of copper}}{63.546 \text{ g of copper}} \times \frac{1 \text{ mole of copper}}{63.546 \text{ g of copper}} = 0.208 \text{ moles of copper}
\]

Divide each mole value by the smallest number of moles (0.208).

\[
\frac{0.6260}{0.208} = 3.01 \quad \frac{0.208}{0.208} = 1.00
\]
There is a ratio of 3 moles of oxalate to one mole of copper. To write the empirical formula of 
$K_xCu_y(C_2O_4)_z \cdot 2H_2O$ you must know the charges of all the ions in the coordination compound. You have $Cu^{2+}$, $C_2O_4^{2-}$ and $K^+$. So far you know $y = 1$ and that $z = 3$ that would mean $x$ would have to equal 4 to give you the neutral compound. So your empirical formula is $K_4Cu(C_2O_4)\cdot 2H_2O$. You are to assume that the empirical formula is also the molecular formula of the coordination compound. You can now calculate the theoretical and percent yields of your compound using the following equation.

\[
y \text{CuSO}_4 \cdot 5\text{H}_2\text{O} + \text{excess reagent} \rightarrow K_xCu_y(C_2O_4)_z \cdot 2\text{H}_2\text{O} + \text{other products}
\]

You are to assume that $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ is the limiting reagent. Balance the copper on either side of this equation and you get

\[
\text{CuSO}_4 \cdot 5\text{H}_2\text{O} + \text{excess reagent} \rightarrow K_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O} + \text{other products}
\]

The mole ratio between $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ and $K_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O}$ is 1:1 since $y$ is equal to 1 so the moles of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ equal the moles of $K_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O}$.

In order to calculate the theoretical yield of the complex salt you must first calculate the moles of starting material $(\text{CuSO}_4 \cdot 5\text{H}_2\text{O})$ from the mass of starting material you weighed out the first week of this laboratory experiment, which can be found from your data under Part I.

\[
\frac{\text{mass of CuSO}_4 \cdot 5\text{H}_2\text{O}}{\text{MW of CuSO}_4 \cdot 5\text{H}_2\text{O}} = \text{moles of CuSO}_4 \cdot 5\text{H}_2\text{O}
\]

Then you must convert moles of starting material to moles of the complex salt $(K_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O})$.

In this example the moles of $K_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O}$ are the same as the moles of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

To then determine the theoretical yield (in grams) of your complex salt, you take the moles of your complex salt (same number of moles as $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$) and convert to grams using the molecular weight of the complex salt.

\[
\text{Theoretical yield of K}_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O} = \text{moles of K}_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O} \times \text{MW of K}_4\text{Cu}(C_2\text{O}_4)_3 \cdot 2\text{H}_2\text{O}
\]

The percent yield is determined from taking the experimental yield (determined before performing part IIB) and the theoretical yield which you calculated.

\[
\% \text{ Yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\%
\]

You should have the following information written in your laboratory notebook:

Empirical Formula of Coordination Compound
Theoretical Yield of Coordination Compound
% Yield of Coordination Compound

Calculations are to be shown in your laboratory notebook.