

**PURDUE UNIVERSITY INSTRUMENT VAN PROJECT**

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**THE SPECTROPHOTOMETRIC DETERMINATION  
OF THE COPPER CONTENT  
IN THE COPPER-CLAD PENNY  
(2-6-96)**

PURPOSE

The purpose of this experiment is to determine: (1) the percentage of copper in a copper-clad penny and (2) the thickness of the copper layer on the copper-clad penny. These analyses will employ spectrophotometric techniques.

BACKGROUND

Pennies minted in the United States since 1982 no longer contain pure copper metal. This change was due to the fact that the cost of the copper metal required to produce a penny was higher than the face value of the penny. In fact, pennies minted after 1982 consist of a copper “coating” on a core that is comprised of an alloy containing both zinc and copper (the core is mostly zinc, however). This is not the first time that zinc has been used in pennies. Perhaps you have seen the “steel-gray” pennies minted in 1943 which were a result of World War II. Copper at that time was being conserved for the war effort and the pennies minted in that year consisted of a zinc “coating” on a steel core. Moreover, due to the high cost of silver, all other coins minted in the United States no longer contain this precious metal. The silver appearance of nickels, dimes, and quarters is due to nickel metal which is used along with copper in these coins. For example, if you look at the edge of a dime or a quarter, you will clearly see a layer of copper!

INTRODUCTION

**Complex ions** are ions formed by the bonding of a metal atom or ion to two or more **ligands** by coordinate covalent bonds. A **ligand** is a negative ion or neutral molecule attached to the central metal ion in a complex ion. Many of these species are highly colored due to their ability to absorb light in the visible region of the electromagnetic spectrum. In this experiment, you will first dissolve a copper-clad penny in a concentrated aqueous solution of nitric acid, HNO<sub>3</sub>. In aqueous solution, most of the first-row transition metals form **octahedral complex ions** with water as their ligands as shown below in Equations 1 and 2:

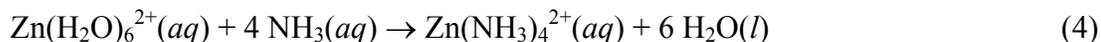
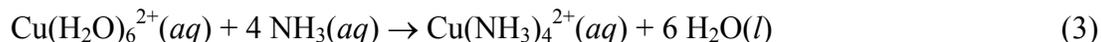


once the penny has been dissolved, you will then convert the aquated copper and zinc complex

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ions to their **tetraamine** complex ions (ie., by replacing the H<sub>2</sub>O ligands with ammonia, NH<sub>3</sub>, ligands) as shown below in Equations 3 and 4:



You can detect the presence of the Cu(NH<sub>3</sub>)<sub>4</sub><sup>2+</sup> ion by its characteristic deep-blue color. Not only can you see the blue color, but you can measure its intensity with a spectrophotometer. By using the spectrophotometer, you will be able to make measurements that will make it possible for you to determine the percentage of copper in a penny.

### PROCEDURE

#### PART I. PHYSICAL PROPERTIES, APPEARANCE, AND DIMENSIONS OF THE PENNY

1. Obtain a penny minted after 1982. Weigh the penny on the analytical balance to the nearest 0.0001 g and record the data in TABLE 1.
2. With a ruler, carefully measure the thickness and the diameter of the penny to the nearest 0.1 mm. Record your data in TABLE 1. Note that we will assume that the penny is perfectly cylindrical (although this is not strictly correct).
3. In TABLE 1, record the year of your penny and the mint that produced it. For example, if there is a small "P" or no letter below the year, the penny was minted in Philadelphia. If there is a small "D" below the year, the penny was minted in Denver.

TABLE 1

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weight of penny	g
thickness of penny	mm
diameter of penny	mm
year minted	
produced by mint in	
appearance of penny	

#### PART II. PREPARING THE PENNY FOR ANALYSIS

- Place your penny in a 100 mL beaker and label the beaker with your group's names.
- Give your beaker to your teacher so that he/she may add the nitric acid. **DO NOT ATTEMPT TO DO THIS ON YOUR OWN!** In the fume hood your teacher will measure out 20 mL of 8 M HNO<sub>3</sub> in a graduated cylinder and add this to the beaker. **NOTE: The reaction you will observe generates NO<sub>2</sub> gas which is highly toxic. Do not allow your face to get to close to the beaker and DO NOT breathe the fumes!** The reaction of the copper and zinc metals in the penny with HNO<sub>3</sub> is quite vigorous so you will not need to stir the reaction.
- Cover your beaker with a watch glass, leave the beaker in the fume hood so that the penny completely dissolves, and go on to the next part of the experiment.

#### PART III. CONSTRUCTION OF A CALIBRATION CURVE FOR Cu(NH<sub>3</sub>)<sub>4</sub><sup>2+</sup>

7. First you will construct a calibration curve that relates the measured absorbance, *A*, to known concentrations of the Cu(NH<sub>3</sub>)<sub>4</sub><sup>2+</sup> ion using the **Beer-Lambert Law**. You will then use the calibration curve to determine the concentration of Cu(NH<sub>3</sub>)<sub>4</sub><sup>2+</sup> in the solution prepared from your penny.

As you have seen previously, concentration and absorbance are related according to the Beer-Lambert Law (Equation 5):

$$A = \epsilon l C \quad (5)$$

where *A* is the absorbance of the species,

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$\epsilon$  is the molar absorptivity (a constant that indicates how well the species absorbs light of a particular wavelength, in units of  $M^{-1} \text{ cm}^{-1}$ ).

$l$  is the path length that the light must travel through the solution (1.00 cm for the cuvet),

$C$  is the concentration (in mol/L).

8. Then, you will plot a graph of measured absorbance versus concentration to determine the molar absorptivity,  $\epsilon$ , for the  $\text{Cu}(\text{NH}_3)_4^{2+}$  ion. Then, you will use this value for  $\epsilon$  to find the concentration of the  $\text{Cu}(\text{NH}_3)_4^{2+}$  ion in an unknown sample (ie., the solution which your penny has been dissolved).

#### PREPARING A STANDARD STOCK SOLUTION OF $\text{Cu}^{2+}$

9. Prepare the standard stock solution as follows. On an analytical balance, weigh out between 0.89-0.93 g of  $\text{Cu}(\text{NO}_3)_2 \cdot 2.5 \text{ H}_2\text{O}$  into a weighing cup and record the weight to the nearest 0.0001 g.

10. Transfer the  $\text{Cu}(\text{NO}_3)_2 \cdot 2.5 \text{ H}_2\text{O}$  to a clean, dry 25 mL volumetric flask. Add distilled water up to the mark on the neck of the flask. Cover the flask with a small square of Parafilm® and shake well to thoroughly mix the solution.

11. You will need a 25 mL buret, buret clamp, and ringstand. Rinse the buret several times with distilled water, letting water flow through the stopcock and tip during each rinse. If the buret does not drain cleanly (ie., if water droplets remain on the sides), wash the buret by using a detergent solution and a buret brush and rinse thoroughly first with tap water and then with distilled water.

12. Rinse and fill the buret with the standard stock solution of  $\text{Cu}(\text{NO}_3)_2 \cdot 2.5 \text{ H}_2\text{O}$  and adjust the level to a suitable reading by letting the solution flow through the stopcock and tip. **BE SURE THAT THERE ARE NO AIR BUBBLES IN THE TIP OR ALONG THE INSIDE SURFACE OF THE BURET.** It is helpful to tip burets slightly and pour fluids down the insides to avoid entrapment of air bubbles. Record your initial buret reading.

13. Clean the 25 mL volumetric flask that you used in PART III (it does not have to be completely dry, but it should be **clean**) and obtain three more 25 mL volumetric flasks. Number the four flasks 1 through 4.

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14. Use the buret to add to each of the numbered volumetric flasks the amounts of the  $\text{Cu}(\text{NO}_3)_2 \cdot 2.5 \text{H}_2\text{O}$  standard stock solution shown in TABLE 2. In addition, using a 10 mL graduated cylinder add about 2.5 mL of distilled water to each of the four volumetric flasks.
15. **IN THE FUME HOOD WITH YOUR TEACHER'S ASSISTANCE**, add 2.0 mL of 15 M  $\text{NH}_3$  carefully from the buret, with swirling, to each volumetric flask until the light-blue precipitate that initially forms dissolves and a deep- blue solution results.
16. Add distilled water to the mark in each volumetric flask, cover each flask with a small square of Parafilm and shake to mix. These solutions are your **standard solutions** for the calibration curve.
17. Obtain two cuvetts. One will be used to calibrate the spectrophotometer the other for the standard solutions.
18. Set the wavelength to 580 nm.
19. Zero and 100% the spectrophotometer using distilled water.
20. Rinse the other cuvet with small portions of the solution in volumetric flask 1 and discard the rinses in your waste container (**KEEP ALL WASTE FOR DISPOSAL AT THE END OF THE EXPERIMENT**). Then, add the solution from volumetric flask 1 to the cuvet until it is about 3/4 full and insert it into the spectrophotometer. Because the % transmittance scale is easier to read on the spectrophotometer, you will record the % transmittance and then convert this value to absorbance. Record the % transmittance of the solution in TABLE 2.
21. Repeat this process for the remaining three solutions (volumetric flasks 2,3, and 4) using the second cuvet and always rinsing the cuvet carefully with the next solution before measuring the % transmittance. **BE SURE TO DISCARD ANY WASTE INTO YOUR WASTE CONTAINER!** After each measurement of your standard solution, check the Zero and 100 % transmittance on your spectrophotometer.
22. You will need the 25 mL volumetric flasks again for PART IV of the experiment. Once you have collected all of the data for the calibration curve, discard your standard solutions in your **WASTE CONTAINER**. Rinse the flasks with small portions of distilled water and discard the rinses in the waste container. Invert the volumetric flasks on a piece of paper towel so that they

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may be dry before you need them again in PART IV.

TABLE 2

flask	volume of $\text{Cu}(\text{NO}_3)_2 \cdot 2.5 \text{H}_2\text{O}$ solution (mL)	% transmittance	absorbance	$\text{Cu}(\text{NH}_3)_4^{2+}$ (mol/L)
1				
2				
3				
4				

PART IV Determination of the copper content in a penny.

23. **IN THE FUME HOOD WITH YOUR TEACHER'S ASSISTANCE**, transfer the solution in which you dissolved your penny to a 100 mL volumetric flask. Use a small portion of distilled water to rinse down the watch glass and the sides of the beaker and transfer the washes to the same 100 mL volumetric flask. Add distilled water to the mark, cover the flask with a small square of Parafilm and shake to mix.

24. Obtain three clean 25 mL volumetric flasks (they do not need to be completely dry, but they must be clean) and label them 5, 6, and 7.

25. Using a 10 mL volumetric pipet, transfer 10.00 mL of the penny solution to each of the three volumetric flasks.

26. **IN THE FUME HOOD**, add 2.0 mL of 15 M  $\text{NH}_3$  carefully from the buret to each volumetric flask until the light-blue precipitate that initially forms dissolves and a deep-blue solution results. Fill each volumetric flask to the mark with distilled water, cover with a small square of Parafilm and shake to mix.

27. Rinse a cuvet with small portions of the solution in volumetric flask 5 and discard the rinses into your **WASTE CONTAINER (NOT DOWN THE DRAIN!)**. Then, add the solution from volumetric flask 5 to the cuvet until it is about 3/4 full and insert it into the spectrophotometer. Record the % transmittance in TABLE 3. Repeat this process for the remaining two solutions (volumetric flasks 6 and 7) always rinsing the cuvet carefully with the next solution before measuring the % transmittance. Periodically check the Zero and 100 % T readings as described

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above and reset if necessary.

**COLLECT ALL REMAINING WASTE SOLUTIONS (STANDARD SOLUTIONS AND PENNY SOLUTIONS). WE WILL DO ANOTHER ACTIVITY AT THE END OF THIS EXPERIMENT THAT WILL PROPERLY DISPOSE OF THE WASTE.**

TABLE 3

flask	% transmittance	absorbance	$\text{Cu}(\text{NH}_3)_4^{2+}$ (mol/L)
5			
6			
7			

### DATA ANALYSIS/CALCULATIONS

1. Calculate the absorbances for each standard solution using the transmittance values obtained in PART III and record your data in TABLE 2. Absorbance and % transmittance are related as below in Equation 6:

$$A = -\log(\%T/100) \quad (6)$$

2. Calculate the concentration of the  $\text{Cu}(\text{NH}_3)_4^{2+}$  ion in each standard solution using the information from PART III and record your data in TABLE 2. Assume that all of the copper initially present in the  $\text{Cu}(\text{NO}_3)_2 \cdot 2.5 \text{H}_2\text{O}$  ends up as  $\text{Cu}(\text{NH}_3)_4^{2+}$  (you will check this assumption later).

3. Graph your data from the standard solutions, plot absorbance as the ordinate (y-axis) versus the concentration of the  $\text{Cu}(\text{NH}_3)_4^{2+}$  ion as the abscissa (x-axis) including your data for each standard solution. Create a least-squares line through your data points and determine the slope of the line. The slope of the line is equal to the molar absorptivity,  $\epsilon$ , in  $\text{M}^{-1} \text{cm}^{-1}$ , of the  $\text{Cu}(\text{NH}_3)_4^{2+}$  ion at a wavelength of 580 nm.

4. Calculate the absorbances for each penny solution prepared using the % transmittance values obtained. Record your data in TABLE 3.

5. Using the calibration curve constructed earlier, calculate the concentration of the  $\text{Cu}(\text{NH}_3)_4^{2+}$  ion in each of your three penny solutions and record your data in TABLE 3. Assume that all of

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the copper initially present in the penny ends up as  $\text{Cu}(\text{NH}_3)_4^{2+}$ .

6. Using the concentrations you have just calculated in #2 and taking into account the various dilutions performed, calculate the mass in grams of copper initially present in your penny for each of the three trials. For each trial, calculate the **percentage** of copper in your penny by dividing the mass of copper in your penny by the total mass of your penny and multiplying this result by 100. Calculate the average mass of copper and the average percentage of copper in your penny from the three trials.

7. In the copper-clad penny, the core contains 0.8% copper and 99.2% zinc by mass. Because we want to calculate the thickness of the copper layer on the core, you will need to subtract the mass of copper that is in the core from the total mass of copper that you determined spectrophotometrically. This will give the mass of copper that is in the copper shell. For example, the total mass of zinc in the penny is equal to the total mass of the penny minus the total mass of the copper in the penny:

$$\text{Total mass of Zn} = \text{Total mass of penny} - \text{Total mass of Cu} \quad (7)$$

Now, all of the zinc that is present in the penny is in the core. Since 99.2% of the total mass of the core is due to zinc, you can use your previously calculated value for the total mass of Zn and calculate the mass of the core:

$$\text{Total mass of Zn} = 0.992 (\text{Total mass of core}) \quad (8)$$

Once you have calculated the total mass of the core, the mass of copper present in the core can be calculated by subtracting the total mass of zinc from the total mass of the core:

$$\text{Mass of Cu in core} = \text{Total mass of core} - \text{Total mass of Zn} \quad (9)$$

The mass of copper in the coating is then simply the total mass of copper minus the mass of copper in the core:

$$\text{Mass of Cu in coating} = \text{Total mass of Cu} - \text{Mass of Cu in core} \quad (10)$$

8. Using your data, calculate the **volume** (in  $\text{cm}^3$ ) of copper present in the copper coating for your penny using the known density of copper. You may use the average values that you calculated for your penny.

9. Using the volume of copper that you calculated for the copper coating, calculate the **thickness** (in cm) of the copper coating on your penny using the dimensions you recorded earlier. This can

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be computed by dividing the volume of the copper coating (in  $\text{cm}^3$ ) by the total surface area of the penny (in  $\text{cm}^2$ ). Recall that we are assuming the penny to be perfectly cylindrical. Thus, the surface area of a cylinder is given by the formula shown below in Equation 11:

$$\pi(d/2)^2 + \pi(d/2)^2 + \pi(d)(t) \quad (11)$$

where **d** is the diameter of the penny (in cm),

**t** is the thickness of the penny (in cm),

$\pi = 3.14159$  (a constant)

10. Collect the following data from the other groups and record the data in TABLE 4. **MAKE SURE THAT THE OTHER GROUPS GET *YOUR* DATA AS WELL:**

1. All of the data in Table 1
2. Average mass of copper in penny.
3. Average percentage of copper in penny.
4. Thickness of copper coating in penny.



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#### RESULTS

Record the minting date of the penny you analyzed with the thickness of the copper layer on the penny and the % by mass of copper in the penny.

#### DISCUSSION/ERROR ANALYSIS

Include a discussion of the following in your lab report:

Compare the information about various pennies that is summarized in TABLE 4. Describe any similarities and/or differences in the total percentage of copper and in the thickness of the copper layer for pennies:

1. Minted in different years,
2. Minted in different cities,
3. With different initial appearances.

#### FOLLOW-UP QUESTIONS

1. When you carefully added ammonia to your solutions containing copper, a light-blue precipitate formed initially that eventually disappeared as more ammonia was added. What do you think this precipitate was? Explain your reasoning.
2. If the atomic radius of a copper atom is  $1.28 \times 10^{-8}$  cm, how many **atoms** thick is the copper coating on your penny?

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**THE SPECTROPHOTOMETRIC DETERMINATION  
OF THE COPPER CONTENT  
IN THE COPPER-CLAD PENNY**

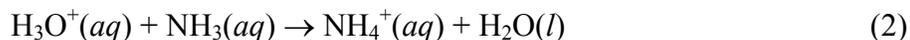
**OPTIONAL WASTE TREATMENT  
(2-6-96)**

INTRODUCTION

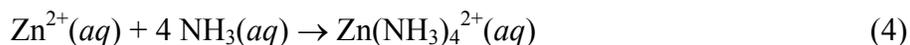
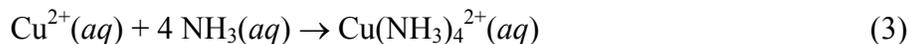
Many waste water treatment facilities have set strict guidelines on the amounts of some transition metals that can be discharged into the local sewer systems. This creates a problem because of some student experiments involved in general chemistry courses can lead to large volumes of aqueous wastes containing high concentrations of some of these transition metals. Such wastes typically have to be collected, taken to a landfill and subsequently buried.

When a penny was analyzed for its copper content, a large amount of aqueous waste that contained high concentrations of zinc and copper was generated. Because ammonia (a weak base) was added to most of the solutions prepared in this experiment, the waste is slightly basic (ie., pH 9-10) and the copper is present as the royal-blue, tetraamine complex ion,  $\text{Cu}(\text{NH}_3)_4^{2+}$ . However, the zinc is probably present as not only the colorless, tetraamine complex ion  $\text{Zn}(\text{NH}_3)_4^{2+}$ , but also as the slightly soluble hydroxide salt,  $\text{Zn}(\text{OH})_2$  ( $K_{\text{sp}} = 4.5 \times 10^{-17}$ ). To remove the zinc and copper from the waste mixture, we will precipitate them as their free  $\text{Cu}^{2+}$  and  $\text{Zn}^{2+}$  ions are present in the solution before the sulfide is added.

Because the zinc and copper tetraamine complex ions are relatively stable species, we will first remove the ammonia from the solution by adding a strong acid. We will use concentrated HCl to react with the weak base, ammonia:



Thus, we can destroy the complex ions of both zinc and copper by adding sufficient acid to shift the following complex ion equilibria completely to the left (LeChatlier's principle):

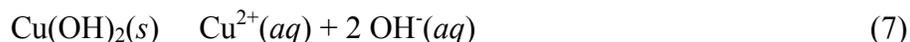
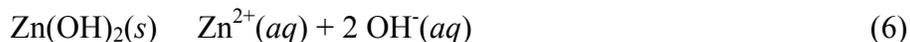
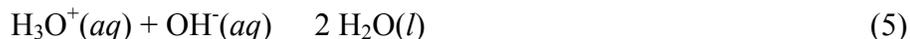


The addition of a sufficient amount of acid will also destroy any zinc hydroxide (or copper hydroxide) in the waste mixture because the reaction of the strong acid,  $\text{H}_3\text{O}^+$ , with the strong

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base, OH<sup>-</sup>, will shift the solubility product equilibria shown below in Equations 6 and 7 completely to the right:



At this point, the waste solution contains the aqueous zinc and copper species, Zn(H<sub>2</sub>O)<sub>6</sub><sup>2+</sup> and Cu(H<sub>2</sub>O)<sub>6</sub><sup>2+</sup>, respectively (recall that the first-row transition metals exist as the aquated, octahedral complex ions in aqueous solution, but we often abbreviate these as Zn<sup>2+</sup>(aq) and Cu<sup>2+</sup>(aq)). Now, if we add a source of sulfide ion, S<sup>2-</sup>, to the waste solution, we should be able to precipitate nearly all of the zinc (K<sub>sp</sub>(ZnS) = 1 x 10<sup>-27</sup>) and copper (K<sub>sp</sub>(CuS) = 6.7 x 10<sup>-42</sup>) without having to worry about the presence of other *soluble* species (ie., such as the complex ions):



We might consider simply filtering the mixture at this point and discharging the filtrate. However, we should first consider the pH of the solution. After we have added sulfide to the solution to precipitate the Cu<sup>2+</sup> and Zn<sup>2+</sup> ions, we would find that the pH of the solution is still slightly acidic. We want to avoid discharging acidic solutions, if possible, so we might want to neutralize (pH = 7) the solution before we discharge it. Second, the ZnS and CuS precipitates are very finely divided and we would have a difficult time trying to filter them directly (they are so fine that some of the precipitate can pass through the filter paper). The solution to the first problem is easy - we can adjust the pH to 7 by adding a small amount of base. The solution to the second problem lies in the ability of some water-soluble **polymers** to act like "flypaper" and "coagulate" very finely divided precipitates into a precipitate that is more easily filtered.

PROCEDURE

**IF ANY REAGENTS ARE SPILLED DURING THIS PROCEDURE, NOTIFY YOUR TEACHER AS SOON AS POSSIBLE!**

PART I. DISSOLVING OH<sup>-</sup> PRECIPITATES AND DESTROYING COMPLEX IONS

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Each group should have about 300 mL of aqueous waste. Divide this waste into two approximately equal portions (ie., 150 mL each) by transferring about half of the waste to a separate 600 mL beaker. Follow the instructions below for each portion.

1. Using a 50 mL graduated cylinder, add 100 mL of deionized water to the waste mixture.

**THE ADDITION OF THE CONCENTRATED HCl AND Na<sub>2</sub>S SOLUTIONS MUST BE PERFORMED IN THE FUME HOOD!**

**KEEP THE CONCENTRATED HCl AND Na<sub>2</sub>S SOLUTIONS IN THE FUME HOOD!**

2. While stirring the waste mixture with a glass stirring rod, add concentrated HCl dropwise (using a disposable pipette) to the waste mixture until the precipitate just dissolves.

#### PART II. PRECIPITATING THE SULFIDES AND ADJUSTING THE pH

3. Using a 50 mL graduated cylinder, add 25 mL of 1.0 M sodium sulfide (Na<sub>2</sub>S) and stir the mixture with the glass stirring rod for 2-3 minutes.
4. Check the pH of the mixture using the pHydrion "A" paper obtained from your teacher. The proper way to check the pH is to dip a glass stirring rod into the solution and then place a drop of the solution on a piece of the pHydrion paper (ie., by touching the end of the stirring rod to the pH paper). **If the pH is equal to or greater than 6, then go on to the next step.** If the pH of the mixture is **less than 6**, add more 1.0 M sodium sulfide (using a disposable pipette) until the pH of the mixture reaches 6. To do this, add the 1.0 M sodium sulfide dropwise (using a disposable pipette), making sure that you **thoroughly stir and properly check the pH** of the mixture after **each addition**. This may take as many as 60 drops.
5. Take the beaker containing the waste mixture back to your bench, add a magnetic stirring bar to the beaker, and stir the waste mixture using a magnetic stirrer. You should stir the mixture well but it should not be vigorous enough to cause splashing.
7. Check the pH of the mixture once again using the pHydrion "A" paper. **If the pH of the mixture is equal to or greater than 8, then go on to the next step.** If the pH is **less than 8**, you will need to adjust the pH of the waste mixture. To do this, add the 6.0 M NaOH dropwise (using a disposable pipette), **making sure that you thoroughly stir and properly check the pH of the mixture after each addition**. This may take as many as 8 drops.

#### PART III. FLOCCULATION OF THE SOLID SULFIDES

8. Make sure that the pH of the waste mixture is equal to or greater than 8. Next, using a plastic

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syringe, add 3 mL of the polymer solution and let the mixture stir for about 5 minutes.

PART IV. SEPARATION OF COMPONENTS IN THE MIXTURE

9. Obtain a 500 mL filter flask, a Buchner funnel and one piece of thick-walled rubber tubing. Connect one end of the tubing to the aspirator on the water faucet and the other end to the sidearm of the clean filter flask. Place a piece of filter paper in the Buchner funnel, turn on the aspirator and wet the filter paper with a small portion of deionized water until the paper completely adheres to the bottom of the Buchner funnel. There should be no gaps around the edges of the filter paper through which the precipitate could pass.

10. Filter the waste mixture through the Buchner funnel. You will need to do this slowly and carefully to ensure that none of the precipitate passes around the edges of the filter paper. If the filtration becomes slow, you may need to remove the filter paper, replace it with a new piece and then continue the filtering process. After you have filtered the bulk of the waste mixture, be sure to transfer any precipitate adhering to the sides of the beaker to the Buchner funnel by rinsing the beaker with a small portion of deionized water.

11. Save the filtrate to test for pH and the presence (or absence) of  $\text{Cu}^{2+}$  and  $\text{Zn}^{2+}$  ions.

12. The filter paper and solid sulfide precipitates are to be put into the waste jar labeled **SOLID and FILTER PAPER WASTE JAR**.

PART V. TESTING THE FILTRATE FOR pH,  $\text{Cu}^{2+}$  and  $\text{Zn}^{2+}$

13. Determine the pH of the filtrate using the pHydrion "A" paper. Record the pH.

14. Test the filtrate for the presence of  $\text{Cu}^{2+}$  and  $\text{Zn}^{2+}$  ions using the following procedures with the commercial test strips that will be available from your teacher.

**$\text{Zn}^{2+}$  Test (to be done by each group)**

**THE ZINC REAGENT IS A VERY CONCENTRATED SOLUTION OF SODIUM HYDROXIDE. IF SKIN CONTACT IS MADE, WASH IMMEDIATELY WITH WATER AND INFORM YOUR TEACHER!**

15. Using a 10 mL graduated cylinder, obtain 5 mL of your group's filtrate and pour the sample into a 100 mL beaker. Add 5 drops of the zinc reagent and stir thoroughly.

16. Dip the  $\text{Zn}^{2+}$  test strip into the sample for about one second. Make sure that the reaction zone

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of the test strip is fully wetted. Remove the test strip, gently shake off the excess liquid and wait for 15 seconds.

17. Compare the reaction zone to the **Zn<sup>2+</sup> Color Scale** found on the container. Record the concentration of Zn<sup>2+</sup> in the filtrate.

**Cu<sup>2+</sup> Test (to be performed by your teacher)**

18. **TO BE DONE BY YOUR TEACHER:** Your teacher will ask for 5 mL of filtrate from one group. If the filtrate has a pH of 8 or higher, your teacher will add 1-2 drops of concentrated sulfuric acid and mix thoroughly. The sample should have a pH between 2 and 6 before testing with the test strip.

19. Dip the Cu<sup>2+</sup> test strip into the sample for about one second. Make sure that the reaction zone of the test strip is fully wetted. Remove the test strip, gently shake off the excess liquid and wait for 30 seconds.

20. Compare the reaction zone to the **Cu<sup>2+</sup> Color Scale** found on the container. Record the concentration of Cu<sup>2+</sup> in the filtrate.

PURDUE UNIVERSITY INSTRUMENT VAN PROJECT

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**TEACHER'S GUIDE**  
**THE SPECTROPHOTOMETRIC DETERMINATION**  
**OF THE COPPER CONTENT**  
**IN THE COPPER-CLAD PENNY**  
(2-6-96)

CLASSROOM USAGE      Chemistry I, ChemCom, Chemistry II, AP Chemistry

CURRICULUM INTEGRATION

Complex ions  
Formation constants  
Solubility products  
Crystal field theory  
Molecular structure of complex ions

REQUIRED KNOWLEDGE

Metric measurement  
Significant digits  
Use of volumetric glassware  
Serial dilutions

PREPARATION

This experiment is labor intensive for the student and the teacher. The teacher may choose to prepare standard solutions to ensure consistent results. The standard solutions are critical for the determination of the copper content of the penny. There are many calculations required in this experiment. They are explained with great detail and should not be a problem.

Areas that have been typed in **BOLD** should be done by the teacher or with the assistance of the teacher. The teacher may choose to dissolve the students' pennies in the fume hood and add the concentrated ammonia in the fume hood to ensure the safety of the students. It is critical that both of these steps be done in a well ventilated area of the room.

TIME              3-4 class periods

SAFETY AND DISPOSAL

## VISIBLE SPECTROSCOPY

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Copper and Zinc compounds should be disposed of according to Flinn disposal 26a

**OR**

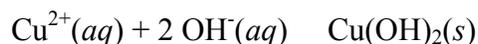
Included with the laboratory is a safe student run disposal of the waster generated by this experiment

### FOLLOW-UP QUESTIONS

1. The precipitate was  $\text{Cu}(\text{NO}_3)_2$ . Aqueous ammonia is a weak base in water and therefore a source of  $\text{OH}^-$ :



The  $K_{sp}$  for  $\text{Cu}(\text{OH})_2 = 5.6 \times 10^{-20}$  so a precipitate forms.



2. Atomic radius:  $r_{\text{Cu}} = 1.28 \times 10^{-8} \text{ cm} \rightarrow d_{\text{Cu}} = 2(1.28 \times 10^{-8} \text{ cm})$

$(9.0 \times 10^{-4} \text{ cm/layer}) \times [1 \text{ Cu atom} / 2(1.28 \times 10^{-8} \text{ cm})] = 35,000 \text{ Cu atoms/layer}$

### ASSESSMENT

Draw a concept map using the following terms:

complex ion	precipitate
ligand	volume
octahedral complex	coating
concentration	thickness
Beer-Lambert Law	core
standards	absorbance
% transmittance	

Vee-diagram of the lab.

Provide the students with an unknown concentration of a copper solution to be determined by

## VISIBLE SPECTROSCOPY

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Group work and division of labor/tasks by group members is strongly encouraged. Each group member as well as their own contributions to the group effort.

LAB WRITTEN BY: Dr. John Nash, Purdue University, Department of Chemistry

ADAPTED BY: Chris Fields

#### REFERENCES

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