LAD A1 Analysis of a Hydrate

Name_____Per____

Introduction

The *polarity* of the water molecule, which makes water a great *solvent* for ionic compounds, causes water molecules to cling to the structure of many solid ionic compounds. When this occurs, the trapped water molecules are called *water of hydration* and they become an integral part of the crystal structure. The resulting compound is called a *hydrate*.

There are many compounds that have a tendency to absorb water vapor from the air. These compounds are said to be *hygroscopic*, and can be used as drying agents. Look for the drying agent in the bottom of the *desiccator* in this lab. You've seen little packets of drying agents in vitamin bottles, leather shoes, and electronics. Some compounds may absorb such large quantities of water vapor that they will actually dissolve in their own water of hydration, a property known as *deliquescence*.

John Dalton was an Englishman, a teacher, and an exceptional theoretical chemist. He developed and wrote the modern atomic theory at the turn of the 19th century (documents suggest he published in 1803). He was influenced by the experiments of two Frenchmen, Antoine Lavoisier and Joseph Louis Proust. A fundamental component of the modern atomic theory is that the mass ratios of elements in a compound will be constant and that the mole ratios of elements in a compound will be *small whole* numbers; this is the *law of constant (definite) composition*, which is sometimes called the *law of definite proportions*. The whole number mole ratio is commonly referred to as the empirical formula of a compound.

One of the challenges in determining the proper chemical formula from information in the periodic table for a compound is that there may be more than one plausible mole ratio for the elements in that compound. This occurs because of the ability of elements to exist in more than one *oxidation state*, particularly many of the *transition metals* and many nonmetals when combining with other nonmetals. Dalton called this the *law of multiple proportions*. For example, if you were testing a compound that contained copper and sulfur, the plausible chemical formula could be CuS or Cu_2S. If you experimentally determine the mass of copper and the mass of sulfur present in a given mass of the compound, you will be able to establish the empirical chemical formula of that particular compound. You could analyze CuS or Cu_2S to verify the law of constant composition; however, you must study both compounds in order to verify the law of multiple proportions.

Procedure Overview

In this *gravimetric* (Google this word for a definition) experiment, you will test an ionic compound containing copper, chlorine, and water molecules locked in the crystal structure of the solid to determine its water of hydration. The general formula for the compound is $Cu_xCl_y \bullet zH_2O$. The letters *x*, *y*, and *z* represent integers that will establish the proper chemical formula for this substance.

This experiment is really two separate procedures. Determining the water of hydration, and establishing the chemical formula for the anhydrous copper(?) chloride.

Although the water molecules are attached to the ionic solid that you will test, they are susceptible to removal by heat, this means they are only loosely attached. In the first part of the experiment, you will **gently** heat a sample of the compound to drive off the water of hydration. By measuring the mass of the sample before and after heating, you can determine the amount of water in the sample and after completing part two, you will be able to calculate its water of hydration, the "*z*" value.

In the second part of this experiment you will conduct a chemical reaction with the copper(?) chloride compound, which will produce elemental copper. By measuring the mass of copper that forms, you will have the necessary information to determine the moles of copper and moles of chlorine in your sample, and you will be able to establish the proper empirical formula, the "*x* and *y*" values.

PreLAD ***This must be done before class, and this page will be turned in. All work must be shown clearly.***

1. Read the Procedure Overview, the Procedure, and Processing the Data sections and then make a Data/Results Table − please do this in Google Spreadsheet. Title the document with your LAST name first, and share it with the teacher with full permissions. All other labs in this course will also be in this very same document under a new tab. Set up your formulas for calculations.

2. Read the Post Lab Questions so you know what is coming and so that you can be discussing the questions with your partners and lab bench neighbors at appropriate times during lab time.

3. The term "gravimetric" was used in the introduction. Look this word up and jot down a concise definition.

4. The concept of *heating to a constant mass* is used in this lab. What does this mean and why is it necessary / important?

5. Beryllium sulfate is a hydrated compound whose formula can be written $BeSO_4 \bullet zH_2O$, where z is the number of moles of H_2O per mole of $BeSO_4$. When a 3.284 g sample of this hydrate was heated to 130°C for 10 minutes, all of the *water of hydration* was lost, leaving 1.948 g of *anhydrous* beryllium sulfate. Calculate the value of z, and write the formula and name of the hydrate.

6. A piece of iron weighing 85.65 g was burned in air. The mass of the iron oxide produced was 118.37 g.
   a. Calculate the moles of iron in the compound.

   b. According to the law of conservation of mass, what is the mass of oxygen that reacted with the iron?

   c. Calculate the number of moles of oxygen in the product.

   d. Use the whole number ratio between the number of moles of iron and number of moles of oxygen to calculate the empirical formula of iron oxide.

   e. Determine the oxidation numbers of the oxygen and iron in this compound. (Yikes! a fractional oxidation number on Fe? Say whaaat? Can you lose a fraction of an electron?) Google *magnetite* and check out Wikipedia to find out what's really going on with the iron ions in this compound. We will discuss this briefly in class − take notes here.

MATERIALS *per lab group*

- tongs, forceps, glass rod, & plastic spatula
- ring & burner
- small beaker
- funnel & paper
- large watch glass
- vial of copper chloride hydrate (~2g)
- aluminum wire (~ 20 cm)
- ~6 M HCl solution in dropping bottle
- tap water wash bottle
- small terra cotta plant pot piece
- lighter or matches
- tile for cooling evaporating dish
- dessicator

On center Lab bench

- evaporating dish and cover (dried in oven)
- balances
- drying ovens (set at 110−120ºC)
- class tray to collect filter paper to dry

PROCEDURE ***Goggles and aprons are not an option - Wear them.***

Part A − Determination of the water of Hydration (Use ~ 2-ish grams of hydrate)

1. Review PreLAD question #5. What data and calculations were necessary to determine the water of hydration? What measurements in the lab you will need to make? *Bullet a procedure below*, and set up your data table accordingly for Part A.
2. Be sure and use the same balance throughout the Lab. Do NOT change midway through. Measure and record the mass of a clean, preheated evaporating dish. (The evaporating dishes were in the hot oven before you arrived and should be slightly warm to the touch.) Be sure and inspect your evaporating dish, observing glazed and unglazed surfaces. (This will allow you to answer a post-LAD question later.) Tare the evaporating dish and then mass something less than 2 g of the copper chloride hydrate into the evaporating dish. Use a glass stirring rod to break up any large pieces of the substance by pressing the pieces against the wall of the evaporating dish AND spread the copper chloride out, rather than in a tall pile − this will make it easier for all of the water to be eliminated.
3. Set up a ring stand, ring, and clay triangle for heating the sample. Make sure the evaporating dish fits comfortably in the ring (clay triangle if necessary). Set up a lab burner and ignite the burner away from the evaporating dish. Adjust the burner to get a small, blue flame.

   Hold the burner in your hand and move the flame slowly back and forth underneath the evaporating dish to GENTLY heat the sample. DO NOT OVERHEAT the compound. Note the color change, from blue-green to brownish, as the water of hydration is driven out of the crystals. When the sample has turned brownish, gently heat the evaporating dish for two more minutes.

   Remove and turn off the burner. Use the tongs as demonstrated in class to remove the evaporating dish from the clay triangle. Place the evaporating dish in the desiccator and allow the sample to cool for 5 minutes or so, until you can touch it. Mass the evaporating dish and anhydrate sample. Be sure and observe the color of the substance in the bottom of the desiccator.
4. Reheat the evaporating dish, and sample until constant mass is achieved. (Within 0.08 g if possible − check with the teacher before doing a third or fourth heating.)

Part B − Determination of the Empirical Formula ***Goggles and aprons are not an option - Wear them.***

Observations: In the space below, comment on the color of the anhydrate, then the color as you begin to dissolve the anhydrate, paying attention to how the color changes as more water is added.

5. Dump your anhydrate into a small beaker. No need to weigh the beaker. Add a small amount of tap water to the evaporating dish, and stir with an stirring rod to completely dissolve the solid that has stuck and pour contents into the small beaker dissolving the anhydrate. Repeat the rinsing process as necessary. (If by chance, the anhydrate is too stuck, you may need to leave the evaporating dish to soak overnight.

Check in to see if this will be a stopping point in the lab, to be continued tomorrow.

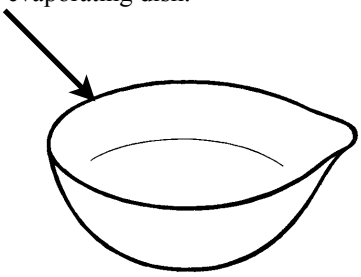
6. Take your piece of wire and curl as demonstrated. The reaction may take a few minutes to begin, then it will proceed quite rapidly. The reaction will take some time to complete. The color of the solution should give some indication as to when the reaction is completed. Explain in the space below.

Observations: In the space below, comment on observations that give evidence that a reaction is occurring.

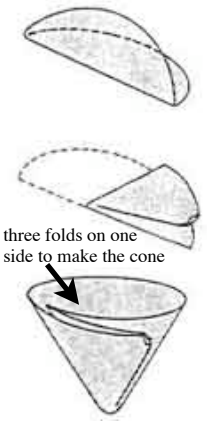
5

7. When the reaction seems completed, shake the wire to remove the copper. Use a plastic spatula and forceps to get as much copper as possible off the Al wire. The left-over wire can be put into the trash can.
8. If the solution is cloudy, check with the teacher − you may need to add a squirt of 6 M HCl solution to dissolve any insoluble aluminum salts in the mixture. The acid should make the solution clear. CAUTION: Handle the hydrochloric acid with care. It can cause painful burns if it comes in contact with the skin.
9. Collect the copper by gravity filtration. Weigh your filter paper, set into funnel and moisten with wash water to hold paper in funnel. You may be able to pour some of the solution off by decanting without losing any of the copper. Pour the remaining solution and copper into the filter paper. As the solution drains through, add some rinse water.
10. Take the filter paper out of the funnel and spread out on to a large watch glass (labeled with your name) and set the watch glass on the class tray to dry overnight.
11. Weigh the dried copper on filter paper.
12. Use a sponge as necessary to clean your lab area, tidy up the tray and leave all other materials on the lab tray at your lab station. Wash your hands when finished cleaning up.

Take a close look at the rim of the evaporating dish.



Folding filter paper



PROCESSING THE DATA

All of the calculations below should be line items in your data/results table. Since # 8 is the grand result of this lab, be sure you have room to report it below your data/results table.

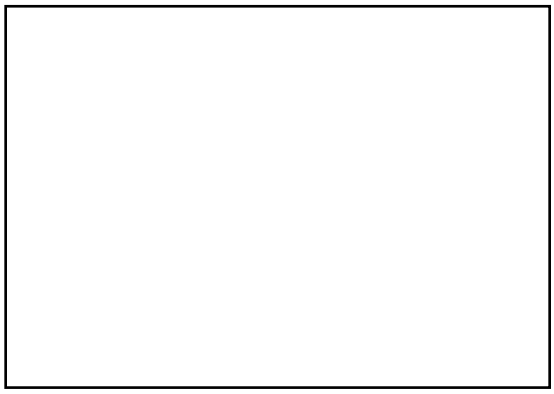
- You (or your partner) must enter your measured data (with partner's name) to the Lab data collection form on the website.
- Be sure that you have a second column in your data/results table with the SAMPLE DATA.

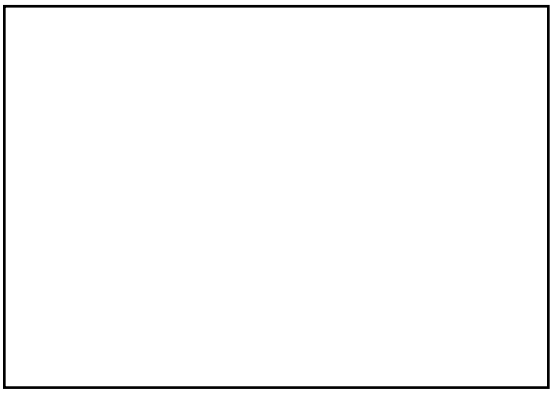
1. Calculate the final constant mass of your anhydrate (by subtracting the empty dish)
2. Calculate the mass of water removed from the hydrate
    - Calculate the number of moles of water were in your sample of copper chloride hydrate.
3. Calculate the mass of copper that was in your sample of copper chloride.
    - Calculate the number of moles of copper that were in your sample of copper chloride.
4. Using the Part B anhydrate, calculate the mass of chlorine were in your part B sample of copper chloride.
    - Calculate the number of moles of chlorine were in your sample of copper chloride.
5. Calculate the Cl/Cu mole ratio. Round to one decimal place.
6. Calculate the moles of anhydrate.
7. Calculate the mole water/mole anhydrate (part A) ratio.
8. Write the proper chemical formula and name for the original hydrated compound that you tested.
9. Using the mass of the **hydrate** calculate the theoretical yield of copper.
10. Calculate the percent yield of copper actually produced in this experiment.

POST LAD QUESTIONS − ***To be completed on this page. Clear, concise, complete answers − full sentences are not necessary.***

1. Now that you know the formula and therefore the charge on the copper in the copper chloride, write a balanced *net ionic* equation to represent the redox reaction between the aluminum metal and the copper chloride solution.

2. During part B of the lab, the copper chloride anhydrate was dissolved in water. Draw a particulate sketch of a single formula unit of dissolved copper chloride in the box on the left. In the box on the right, sketch at least four or five water molecules as they would arrange around a single dissolved copper chloride unit. Indicate the partial charges of the water molecules by marking with the symbols: δ+ or δ− to indicate the partial polar charges.
   In the space below each drawing, make a bullet list of a several key points that each model attempts to display.





3. The arrangement of particles demonstrated in the box on the right is called a *hydration shell*. This causes the copper metal becomes a *complex ion* with its surrounding water molecules. For metals other than group I and II, some water molecules are ripped apart due to their close association with the ions resulting in some H^+ ions ending up in solution. This process can be expressed in the reaction shown below and can be called *hydrolysis*. What does the double arrow shown in the reaction tell us about the completion of the reaction, and what is actually found in the beaker: reactants, products, or both?

$$Cu(H_2O)_6^{2+} \rightleftarrows [Cu(H_2O)_5(OH)]^+ + H^+$$

4. Recall that you may have seen bubbles of gas forming during the redox reaction. These bubbles are due to a side reaction between some of the aluminum wire and acid (H^+ ions) formed by the hydrolysis of the copper ions. Write a balanced *net ionic* reaction that occurs between aluminum metal and acid ions. This is a redox reaction. Which element is oxidized, which element is reduced.

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5. Notice that the rim of the porcelain evaporating dish is unglazed. Observe what water will do to the unglazed terra cotta plant pot on your tray. This is why we preheated the evaporating dish prior to measuring the evaporating dishes initial mass. Would the experimentally determined moles of water in the hydrate appear to be larger, smaller, or the same on a humid day if the evaporating dish had not been pre-heated before the initial weighing? Justify your answer.

6. Why should objects be cooled before their mass is determined on a sensitive balance? (More than one answer would be appropriate, but do NOT answer that the hot dish will not break the balance, because in fact a hot dish would probably NOT break or damage the balance.)

7. After removing the water from the hydrated sample a student fails to cool the evaporating dish in the desiccator. This might be a problem on a humid day in September since our lab is not air conditioned. Would this oversight cause the experimentally determined mole ratio of the water /anhydrate to be larger, smaller, or remain the same. Justify your answer.

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Scoring Rubric (out of 100 pts)

5 PreLab Data Table done BEFORE Lab day, including embedded formulas as best you can.

- The expectation is that the data table is set up in your Google Lab Sheet the day before we begin the lab. This means before *10 pm the night before* we will begin the lab. These points can not be recovered if the work is not done ahead of time.
- The expectation is that the prelab questions will be done *before* class begins on the day for which they have been assigned.

35 Questions

- These points are distributed among the PreLab and PostLab questions. Work must be shown for any calculations. While you are allowed and encouraged to collaborate with classmates, the expectation is that the answers are not copied from another classmates work.
- Doing the questions before they are due will allow you to ask for help from the teacher as necessary. Do this via email (before the night before they are due) to get on the class-time agenda and or come in after school.

35 Data Table

- The data table should not be simply copied from a classmates who shares their sheet with you. You are expected to make your own. Certainly you can collaborate, but simply cutting and pasting a classmates is not considered collaboration.
- All calculations should be completed with embedded formulas.

25 Post Lab Quiz

- Up to 25 more points are available by successful completion of the Post Lab quiz which will take place during class time after the lab "report" has been graded and returned.

Scoring Rubric	
5	**PRE**
35	Questions
35	Data Table
_____	*Out of 75*
25	Lab Quiz