**Lab 8C: Determining the Formula of a Hydrate Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_Pd\_\_\_\_\_\_\_**

**Objectives**

The purpose of this experiment is to determine the empirical formula of an unknown hydrate. **Hydrates** are compounds which contain a specific number of water molecules loosely attached within the crystal structure. An example is sodium carbonate decahydrate (washing soda). The formula for this substance is Na2CO3•10H2O. Do you recognize sodium carbonate (Na2CO3)? Inside the hydrated form of sodium carbonate are 10 trapped water molecules. Those water molecules are not part of the sodium carbonate compound itself – together they form a hydrate.

Hydrates can normally be broken down into the **anhydrous** compound and water by gentle heating according to the equation below. From the data collected, the number of moles of water will be determined per mole of anhydrous compound to give the empirical formula of your hydrate.

Na2CO3•10H2O 🡪 Na2CO3 + 10H2O

1 mol of Hydrate 🡪 1 mole of Anhydrous compound + 10 moles of water

The water within the hydrate is driven off as steam by heat, leaving behind the **anhydrous** compound (a substance without water). By determining the mass of the hydrate and the mass of the anhydrous compound, you can determine the mass of water that was driven off. Using the mass, you can find the moles of water and the moles of anhydrous compound and the mole ratio between them.

**Prelab Assignment** – Read over the entire lab and answer the following questions on a separate sheet of paper to turn in at the beginning of class.

1. Define anhydrous (or anhydrous compound).

2. How can you turn a hydrate into an anhydrous compound?

3. What evidence should you observe if you are heating a hydrate?

4. You will be using copper (II) sulfate. What is the chemical formula of copper (II) sulfate?

**Materials**
 Bunsen burner Mortar and pestle

 Crucible and lid Crucible tongs

 Copper (II) sulfate Droppers for water at the end

**Determining the empirical formula of a copper (II) sulfate hydrate.**

1. Thoroughly wipe a crucible and cover with a clean paper towel to remove dirt and other particulate matter. Then determine the mass of the crucible (and lid).

2. Obtain about 2 g (about 1 teaspoon) of copper (II) sulfate hydrate sample and transfer the sample to the crucible.

3. Determine the mass of the crucible, hydrate sample, and crucible cover to appropriate sig figs.

4. Place the crucible, lids and contents on the wire ring above the burner or hot plate. Partially cover the opening of the crucible with the cover. Be careful; crucible lids like to slide off and break.

5. Heat your crucible and its contents for at least 10 minutes. You should notice a color change as the water leaves the compound.

6. Using crucible tongs, remove the crucible and lid from the ring and place on the lab bench. As you remove it, set the lid upside down to prevent loss of anhydrous compound. Once it is all on the bench, completely cover the crucible with the lid to prevent moisture from entering and adding mass.

7. Allow the crucible to cool to room temperature. (Hold your hand about 1 cm above the crucible to test.) Then determine the mass of the crucible and contents (and lid) appropriate sig figs.

8. After you have finished obtaining the mass, add a few drops of water to the dehydrated material and observe.

**Data**

1. Mass of empty, dry crucible and lid\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

2. Mass of crucible, lid and hydrate\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 3. Mass of hydrate alone \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

4. Final mass of crucible, lid and anhydrous copper (II) sulfate\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 5. Mass of anhydrous compound alone\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

6. Observations of anhydrous compound after you add a few drops of water:

**Lab Report**

1. Determine the mass of water lost during heating.

 \_\_\_\_\_\_\_\_\_\_\_g water lost

1. Using the mass of water lost, determine the number of moles of water that were lost during heating.

 \_\_\_\_\_\_\_\_\_\_\_\_mol water

1. Determine the molar mass of copper (II) sulfate – recall that you wrote down the formula as part of your prelab.

 \_\_\_\_\_\_\_\_\_\_\_\_g/mol

1. Staring with the mass of anhydrous compound alone (data table #5), calculate the number of moles of anhydrous copper (II) sulfate present in your crucible.

 \_\_\_\_\_\_\_\_\_\_\_\_mol anhydrous

1. Determine the empirical formula of your hydrate by setting up the ratio$\frac{moles of water}{moles anhydrous}$. Round this number to the nearest whole #.

 Empirical formula of copper (II) sulfate hydrate\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_•\_\_\_\_\_\_H2O\_\_\_

1. In step 6, you are told to completely cover the crucible. Why would that be an important step?

7. Determine the empirical formula of the following hydrate of MnSO4 using the lab data below. Show all work, sig figs and units. Follow the same steps as in the lab (find grams water, moles water, moles anhydrous and then the ratio)

* 1. Mass of empty crucible and lid: 75.300g
	2. Mass of crucible, lid and hydrated MnSO4• xH2O: 85.960g
	3. Mass of crucible, lid and anhydrous MnSO4: 83.152g

 MnSO4 •\_\_\_\_\_\_H2O

BVSD Standards 1.7.f: Students can recognize and apply a variety of empirical methods for determining molar mass.; 1.7.e:Students can calculate the empirical and molecular formula of a substance from experimental data.