

Name: _____
Partners: _____
Hour: _____ Date: _____

Chemistry: Lab – Formula of a Hydrate

Introduction:

Many salts that have been crystallized from water solutions appear to be perfectly dry, yet when heated yield large quantities of water. The crystals change form, and sometimes color, as the water is driven off. This suggests that water was present as part of the crystal structure. Such compounds are called **hydrates**. A hydrate that has lost its water is called an **anhydrous salt**. For a hydrate, the number of moles of water present per mole of salt is usually some simple, whole number.

Because salts consist of cations and anions bonded together (and also because all metals are cations and all nonmetals are anions), an anhydrous salt is often symbolized **MN**, where the M stands for “metal” and the N stands for “nonmetal.” Similarly, a hydrate – which consists of an anhydrous salt and water – is often symbolized **MN · ? H₂O**, where the question mark indicates the integer number of water molecules for each formula unit of salt. The dot between the MN and the ? H₂O means that the water molecules are rather loosely attached to the anhydrous salt. When referring to an unknown hydrate, you should use the notation described above.

One example of a hydrate is copper (II) sulfate pentahydrate. Its blue crystals look and feel dry, but each mole of the anhydrous salt is actually bonded to five moles of water. The compound’s formula is **CuSO₄ · 5 H₂O**. The molar mass of CuSO₄ · 5 H₂O is:

$$63.5 \text{ g} + 32.1 \text{ g} + 4 (16.0 \text{ g}) + [5 (18.0 \text{ g})] = 249.6 \text{ g}$$

If a 249.6 g sample of CuSO₄ · 5 H₂O were heated to drive off all the water, the anhydrous salt CuSO₄ would weigh $63.5 \text{ g} + 32.1 \text{ g} + 4 (16.0 \text{ g}) = 159.6 \text{ g}$, which is the mass of one mole of CuSO₄. The mass of water that has been boiled off into the air is $[5 (18.0 \text{ g})] = 90.0 \text{ g}$, which is the mass of five moles of water. The formula of the hydrate shows the ratio of the moles of anhydrous salt to the moles of water; in the above case, that ratio is 1:5.

In this experiment, you will be given a sample of hydrate. You will determine the mass of the water driven off by heating, as well as the amount of anhydrous salt that remains behind. Then, given the mass of one mole of the anhydrous salt, you will determine the empirical formula of the hydrate.

Pre-Laboratory Assignment

Use the information to answer the questions. Show work, include units, and put your answers in the blanks.

William weighs an empty beaker and finds it to have a mass of 95.83 g. After putting a spoonful of an unknown hydrate into the beaker, he finds that the mass has increased slightly to 99.87 g. He heats the beaker and its contents twice, and finds that the mass has dropped to 97.22 g. William is told by his teacher that the molar mass of the anhydrous salt is 74.10 g.

1. What mass of hydrate did William start with?

2. How much water was driven off from the hydrate during the heating process in units of...

A. **grams?**

B. **moles?**

3. How much anhydrous salt remained in the beaker in units of...

A. **grams?**

B. **moles?**

4. Write down the mole ratio as decimal numbers: _____ moles anhydrous salt : _____ moles water

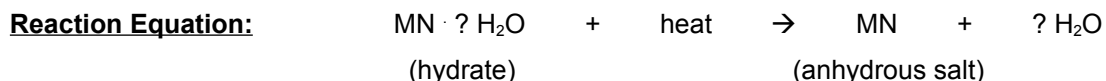
Write down the mole ratio as whole numbers: _____ moles anhydrous salt : _____ moles water

5. What is the formula of the hydrate? (use MN to symbolize the anhydrous salt)

6. Based on William's data, calculate the percentage of water in the sample of hydrate.

7. Why must William heat the sample twice instead of just once?

Laboratory Procedures



Materials: safety glasses hydrate crystals matches
150 mL beaker ring stand with ring bunsen burner
watch glass wire gauze tongs

Safety: Wear safety glasses at all times.
Do not touch a hot beaker with your hands.
Do not weigh a beaker that is still piping hot.

Procedure:

1. Set up a ring stand. Secure the ring on the stand and place the wire gauze on the ring. Place a clean, dry 150 mL beaker on the wire gauze.
2. With the bunsen burner on low heat, warm the beaker for two minutes.
3. Allow the beaker to cool for three minutes, then use the tongs to carry the beaker over to a balance. Weigh the beaker and record this mass in the Data Table.
4. Without using the tare button and while the beaker is still on the balance, place about $\frac{2}{3}$ of a teaspoon of the unknown hydrate crystals into the beaker. Record this mass in the Data Table.
5. Place the beaker back on the wire gauze and heat with the bunsen burner for two minutes on low heat, then for five minutes on high heat.
6. Allow the beaker to cool for five minutes, then use the tongs to carry it back to the same balance you used before. Record the mass in the Data Table.
7. Repeat steps 5 and 6, but only heat the beaker for four minutes this time instead of seven. Record the mass in the Data Table.
8. Your teacher will tell you the molar mass of the anhydrous salt. Record this value in the Data Table.
9. Clean up as directed by your teacher.

Data Table: (Include units, and record all quantities to the nearest 0.01 g.)

<i>Quantity Measured</i>	<i>Mass</i>
dry beaker	
beaker and contents before heating	
beaker and contents after first heating	
beaker and contents after second heating	
molar mass of anhydrous salt (from teacher)	
percentage of water in hydrate (from teacher)	

Processing Your Lab Data

Show your work, include units, and write your answers in the blanks on the right.

1. What mass of hydrate did you start with?

2. How much water was driven off from the hydrate during the heating process in units of...

A. **grams?**

B. **moles?**

3. How much anhydrous salt remained in the beaker in units of...

A. **grams?**

B. **moles?**

4. Write down the mole ratio as decimal numbers: _____ moles anhydrous salt : _____ moles water

Write down the mole ratio as whole numbers: _____ moles anhydrous salt : _____ moles water

5. What was the formula of your hydrate?

6. Based on your data, calculate the percentage of water in the sample of hydrate.

7. Calculate the percent error in your experiment by comparing the actual percentage of water with the percentage you obtained in your experiment using the formula below. Show your work.

$$\% \text{ error} = \left| \frac{\text{Actual} - \text{Experimental}}{\text{Actual}} \right|$$
