Empirical 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 $MN \cdot xH_2O$, where the question mark indicates the integer number of water molecules for each formula unit of salt. The dot between the MN and the x H_2O 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 $\mathbf{CuSO_4 \cdot 5\ H_2O}$. The molar mass of $CuSO_4 \cdot 5\ H_2O$ 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 $CuSO_4 \cdot 5\ H_2O$ were heated to drive off all the water, the anhydrous salt $CuSO_4$ would weigh 63.5 g + 32.1 g + 4 (16.0 g) = 159.6 g, which is the mass of one mole of $CuSO_4$. The mass of water that has been boiled off into the air is [5 (18.0 g)] = 90.0 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. An empirical formula of a chemical compound is the ratio of atoms in simplest whole-number terms of each present element in the compound. For example, Glucose is $C_6H_{12}O_6$; it's empirical formula is CH_2O.

A hydrate is a compound that is chemically combined with water molecules. In contrast, an anhydrate does not contain water, and has had all of its water removed.

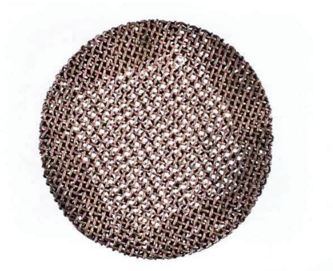
Objective:

1. To identify the empirical formula for an unknown hydrate, selected by your teacher.

Apparatus:

Crucible Hot plate Scale


Wire Gauze BeakerTongs Scoopula



Materials:

$MN \cdot xH_2O$ crystals (an unknown hydrate crystals)

Procedure: (to be flowcharted)

1. First, preheat your crucible on a hot plate for two to three minutes. Your dish should be clean and dry before use.
2. After your crucible is heated, let it cool for at least three minutes. Use your tongs to move the dish to a suitable cool-down location.
3. Once the crucible has cooled down enough for you to touch it without getting burned, place it on a scale and determine the mass to the nearest 0.01g.
4. Place the empty crucible on the scale and record its mass.
5. Add about 2 grams of the provided hydrate .
6. Be sure to record the mass of the crucible and the hydrate to the nearest 0.01g.
7. Place the filled crucible on the hot plate and heat gently until the water has been driven off (Heat until there is no more popping or spattering). Allow the dish to cool.
8. When the crucible is cool enough to touch, transfer it to a scale and find the mass.
9. To make sure all the water is driven off, heat the dish another time (Repeat step 7 & 8). If the results do not agree within 0.02g, consider repeating these steps again

Data & Observations:

Quantity Measured	***Mass***
dry crucible	
crucible and contents before heating	
crucible and contents after first heating	
crucible and contents after second heating	
molar mass of anhydrous salt (from teacher)	
mass of water given off	

Analysis:

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.

   ____

Discussion:

1. Compare and contrast the terms hydrate with anydrate.
2. A 2.815 g sample of $CuSO_4 \bullet xH_2O$ was heated until all of the water was removed. Calculate the percentage of water of hydration and the formula of the hydrate if the residue after heating weighed 2.485 g.
3. The mass percent of water in a hydrate of $MnCl_2$ is 36.41%. What is the empirical formula of the hydrate?

Sources of Error:

List the error in the equipment that was used for QUANTITATIVE DATA only!

Conclusion:

Be sure to make a connection to your everyday life!