**Determining the formula of a hydrated salt**

INTRODUCTION. Some chemical compounds, especially inorganic salts, incorporate water into their crystalline structures. Water has a polar structure: it has positively and negatively charged parts within each molecule. This gives it a strong attraction toward ions. The ions in some salts attract and form strong bonds with water molecules.

These salts, when they have absorbed water, are called hydrates. Anhydrous salts are salts that can form hydrates but which have had all the water driven off, usually by heat. Hydrated salts are characterized by the number of moles of water molecules per mole of salt. The so-called water of hydration of nickel(II) chloride (NiCl2) is six moles H2O for every one mole of NiCl2. The hydrate is called nickel(II) chloride hexahydrate. NiCl2·6H2O The formula of this hydrate shows the molar amount of water incorporated into the crystal matrix.

For most hydrates the amount of water included in the formula is only important when trying to measure molar amounts of the salt. You need to know the true molar mass in order to measure out the mass needed to give a certain number of moles. The chemical importance of the water of hydration is minimal since it can be driven off by heat or simply dissolve away if the salt is dissolved in water. From the compound above, nickel(II) chloride hexahydrate (NiCl2•6H2O) the molar mass is 237.69 g/mol not 129.60 g/mol. The figure 129.60 g/mol is the molar mass of the anhydrous salt.

Greek Prefixes are used to indicate the molar amount of water in hydrated salt.

**1 (mono-) 2 (di-) 3 (tri-) 4 (tetra-) 5 (penta-) 6 (hexa-) 7 (hepta-) 8 (octa-) 9 (nona-) 10 (deca-)**

Formulas for hydrates are written using a dot convention. A dot is used to separate the formula of the salt from the formula of the water of hydration. A numerical coefficient gives the molar amount of water included in the hydrate. Hydrates are named using prefixes for the word hydrate. One key point, the dot is not a multiplication sign. When calculating the molar mass you add the molar mass of water (multiplied by the coefficient).

An everyday example of hydration is concrete. Concrete is made by mixing Portland cement with water and aggregate materials. The aggregate materials are the gravel and sand that add strength to the final concrete. The Portland cement is a mixture of calcium silicates, calcium aluminate, calcium aluminoferrite and gypsum. All of these chemicals absorb water by hydration. This means that concrete does not ‘dry’ in a conventional sense. Instead the water mixed with the concrete combines chemically with the materials in the cement and the resulting hydrates form a strong matrix that holds the concrete together and makes it strong.

Another interesting example of the value of hydration is the incorporation of hydrated building materials (such as concrete, gypsum wall board and plaster). The building materials will not rise above the 100°C boiling point of water until all of the water of hydration has been driven off. This can help keep damage to a minimum until the fire can be put out. In the construction business this is known as passive fire protection.

In this experiment, you will be instructed to determine the mass of a sample of an unknown hydrate by "difference", using a pre-weighed beaker as the container. The substance will be "dehydrated" by heat and weighed again. The loss of mass represents the mass of water in the original sample, which may be expressed as percentage of water of hydration.

To find the formula of the original hydrate, you will determine the molar ratio of anhydrous salt to water, as in an empirical formula calculation. To obtain the required data you must heat the substance until all of the water is driven off.

**Procedure**

1. Determine and record the mass of an empty beaker.
2. Add approximately 3g of hydrated copper (ii) sulfate to the beaker and record the exact amount below.
3. Hold the beaker with tongs and place it on a hotplate.
4. Heat the beaker on medium heat while stirring for about 10 minutes.
5. Continue heating until all water has been removed from the hydrate.
6. Record your observation.
7. Dispose of the anhydrous salt into the waste beaker at the front of the class.

**Data and analysis**

1. Mass of empty beaker \_\_\_\_\_\_\_\_\_\_\_\_ g

2. Mass of beaker + hydrated salt \_\_\_\_\_\_\_\_\_\_\_\_ g

3. Mass of hydrated salt: [2] - [1] \_\_\_\_\_\_\_\_\_\_\_\_ g

4. Mass of beaker + anhydrous salt (after heating) \_\_\_\_\_\_\_\_\_\_\_\_ g

**Calculate the mass percentage of water and the empirical formula of the hydrate.**

5. Mass of water lost by hydrated salt: [2] - [4] \_\_\_\_\_\_\_\_\_\_\_\_ g

6. Mass percentage of water in hydrated salt \_\_\_\_\_\_\_\_\_\_\_\_ % (show calculation)

7. Molar mass of anhydrous salt \_\_\_\_\_\_\_\_\_\_\_\_ g/mol

8. Moles of anhydrous salt \_\_\_\_\_\_\_\_\_\_\_\_ mol

9. Molar mass of water \_\_\_\_\_\_\_\_\_\_\_\_ g/mol

10. Moles of water in hydrated salt: [5] / [9] \_\_\_\_\_\_\_\_\_\_\_\_ mol

11. Ratio of moles of water to moles of anhydrous salt: [10] / [8] \_\_\_\_\_\_\_\_\_\_ ratio

12. Empirical formula of hydrated salt \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (hydrated salt formula)