Chemistry Lecture #57: Formula for a Hydrate

There are some compounds that look like dry crystals, but actually have water molecules trapped inside the crystals. These compounds are called *hydrates*. A hydrate is a compound that has a specific number of water molecules bound to its atoms.

For example, Epsom salt is the common name of a magnesium sulfate compound that contains water molecules. Below is a picture of Epsom salt crystals.

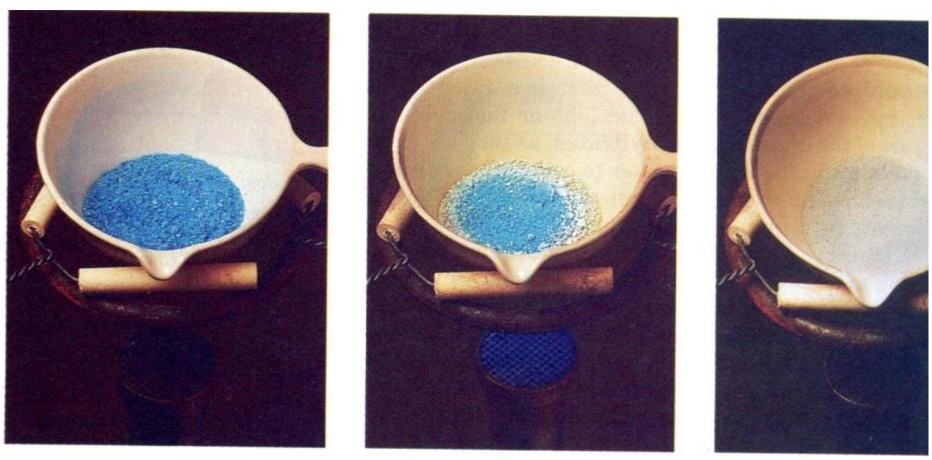
For every formula unit of $MgSO_4$, there are seven water molecules trapped within the crystal. To indicate that there are seven water molecules attached to the magnesium sulfate, we write $MgSO_4$ and tack on $7H_2O$. We separate $MgSO_4$ and $7H_2O$ with a dot. The formula looks like this:

$$MgSO_4 \cdot 7H_2O$$

This compound would be called magnesium sulfate heptahydrate.

Hydrates composed of ionic compounds and water can be heated to remove the water and leave behind the ionic compound as a residue. The residue weighs less than the original hydrate. The difference in mass allows us to calculate the amount of water attached to each formula unit of the ionic compound.

For example, copper (II) sulfate exists as a hydrate. This is a deep blue crystal. But when heated, the water evaporates, and the compound turns white with a faint blue color. Below are pictures showing the change in color that occurs when the hydrate of copper (II) sulfate is heated.



The steps used to calculate the formula of the hydrate after being heated are as follows:

1. Find the mass of water evaporated and the mass of the remaining residue. Convert these values to moles.
2. Divide by the smallest number of moles to get the ratio of Ionic compound residue to water.
3. Write the formula of the hydrate with the number of water molecules in front of H_2O. Put a dot between the ionic compound formula and the number of water molecules.

2.49 g of blue, hydrated copper sulfate is heated. After heating, 1.59 g of white, anhydrous $CuSO_4$ remains. What is the formula for the hydrate? What is the name of the hydrate?

We first need to find the mass of one mole of $CuSO_4$ and one mole of H_2O. You should memorize that 1 mole of H_2O = 18.0 g.

Cu: 1 x 63.5 = 63.5
S: 1 x 32.1 = 32.1
O: 4 x 16.0 = 64.0
160 g/mole (rounded)

H: 2 x 1.01 = 2.02
O: 1 x 16.0 = 16.0
18.0 g/mole (rounded)

Mass of hydrate	=	2.49 g
Mass of $CuSO_4$	=	- 1.59 g
Mass of H_2O	=	0.90 g

$$\frac{1.59 \text{ g } CuSO_4}{1} \times \frac{\text{mole } CuSO_4}{160 \text{ g } CuSO_4} = 0.00994 \text{ mole } CuSO_4$$

$$\frac{0.90 \text{ g } H_2O}{1} \times \frac{\text{mole } H_2O}{18.0 \text{ g } H_2O} = 0.050 \text{ mole } H_2O$$

$$\frac{0.00994 \text{ mole } CuSO_4}{0.00994} = 1 \qquad \frac{0.050 \text{ mole } H_2O}{0.00994} = 5.0 \text{ (2 sig. fig.)}$$

The formula of the hydrate is

$CuSO_4 \cdot 5H_2O$

and is called copper (II) sulfate pentahydrate.

10.41 g of hydrated BaI_2 is heated. After heating, the dry sample has a mass of 9.52 g. What is the formula and name of the hydrate?

Ba: 1 x 137 = 137
I: 2 x 127 = 254
391 g/mole for BaI_2

1 mole H_2O = 18.0 g

Mass of hydrate = 10.41 g
Mass of BaI_2 = - 9.52 g
Mass of H_2O 0.89 g

$$\frac{9.52\text{ g }BaI_2}{1} \times \frac{\text{mole }BaI_2}{391\text{ g }BaI_2} = 0.0243\text{ moles }BaI_2$$

$$\frac{0.89\text{ g }H_2O}{1} \times \frac{\text{mole }H_2O}{18.0\text{ g }H_2O} = 0.049\text{ moles }H_2O\text{ (2 sig. fig.)}$$

$$\frac{0.0243\text{ moles }BaI_2}{0.0243} = 1 \qquad \frac{0.049\text{ moles }H_2O}{0.0243} = 2.0$$

The formula of the hydrate is

$$BaI_2 \cdot 2H_2O$$

and the name of the hydrate is barium iodide dihydrate.