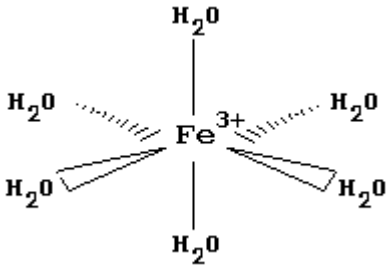
CHEM 109
Introduction to Chemistry
Revision 1.2

Laboratory Exercise: The Chemical Formula of a Hydrate

In this laboratory exercise we will determine the relative amount of Water in a Hydrated compound. We will do this by determining the mass of Water driven out of the compound when it is heated. We can then use this mass determination to find the number of moles Water contained in a sample of the compound.

Hydrates are salts that crystallize from a Water solution and contain weakly bound Water molecules. For example, Iron (III) Chloride, $FeCl_3$, crystallizes with six Water molecules weakly bound to the Fe^{3+} ion in an octahedral arrangement:



Iron (III) Chloride Hexahydrate

This is denoted as $FeCl_3 \cdot 6H_2O$ and is referred to as Iron (III) Chloride Hexahydrate.

Note the unusual method for writing the chemical formula of this compound. A hydrate is an addition compound; a compound that contains two or more simpler compounds. In the above example, the simpler compounds are $FeCl_3$ and H_2O, which combine in a ratio of six Waters to every formula unit of Iron (III) Chloride. This is a weak chemical combination between the Water and the salt. As such, the combination is denoted with a "dot." **However, the Water molecules are as much a part of the compound as the other atoms**.

As another example, Gypsum occurs as a Dihydrate of Calcium Sulfate; $CaSO_4 \cdot 2H_2O$.

When a hydrate is heated, the loosely held water is driven off as Water vapor, leaving an anhydrous salt behind. For example:

$$FeCl_3 \cdot 6H_2O(s) \longrightarrow FeCl_3(s) + 6\ H_2O(g)$$

This dehydration may actually occur in several steps, with the solid crystal rearranging to accommodate the loss of the Water molecules.

The dehydration process is reversible; some anhydrous salts will pick-up significant amounts of Water from the air. Compounds that are particularly prone to picking-up Water from air are referred to as hygroscopic. An example is $CaCl_2$, which picks-up atmospheric water to form the Mono, Di, Tetra and Hexa hydrates.

In our case, we will determine the number of Water molecules per formula unit of Epsom Salt; a hydrate of Magnesium Sulfate:

$MgSO_4 \bullet xH_2O$

Our goal is to determine the number of hydrating Water molecules *x*. We will do this by measuring the mass of a sample of Epsom Salt and then heating it to drive off the Water. The resulting anhydrous sample of Magnesium Sulfate ($MgSO_4$) can be subsequently weighed. The difference in these masses, by the Law of Conservation of Mass, must be the mass of Water driven off. The number of moles Water contained in the sample can then be determined by dividing the mass of the Water by its Molar Mass:

$$\#\ \text{moles}\ H_2O\ =\ \text{mass}\ H_2O\ /\ \text{Molar Mass}\ H_2O$$

Similarly, the number of moles $MgSO_4$ can be determined by dividing the mass of the anhydrous salt by its Molar Mass:

$$\#\ \text{moles}\ MgSO_4\ =\ \text{mass}\ MgSO_4\ /\ \text{Molar Mass}\ MgSO_4$$

These mole numbers are simply a count of the number of Water molecules and Magnesium Sulfate formula units in our sample. The relative number of Water molecules in the Hydrate is simply the count of Water molecules relative to the count Magnesium Sulfate formula units.

$$x\ =\ \#\ \text{molecules}\ H_2O\ /\ \#\ \text{formula units}\ MgSO_4$$

This ratio is equivalent to the ratio of the mole numbers:

$$x = \#\text{ mole } H_2O \,/\, \#\text{ mole } MgSO_4$$

Thus, straightforward mass measurements can be used to determine the relative count of Water molecules in a Hydrated compound.

Procedure

Preliminary Observations

To better understand what is happening during the heating process, we will initially simply observe the dehydration of our Epsom Salt.

1. Obtain approximately 0.5g of Epsom Salt. Note the color and other observations.

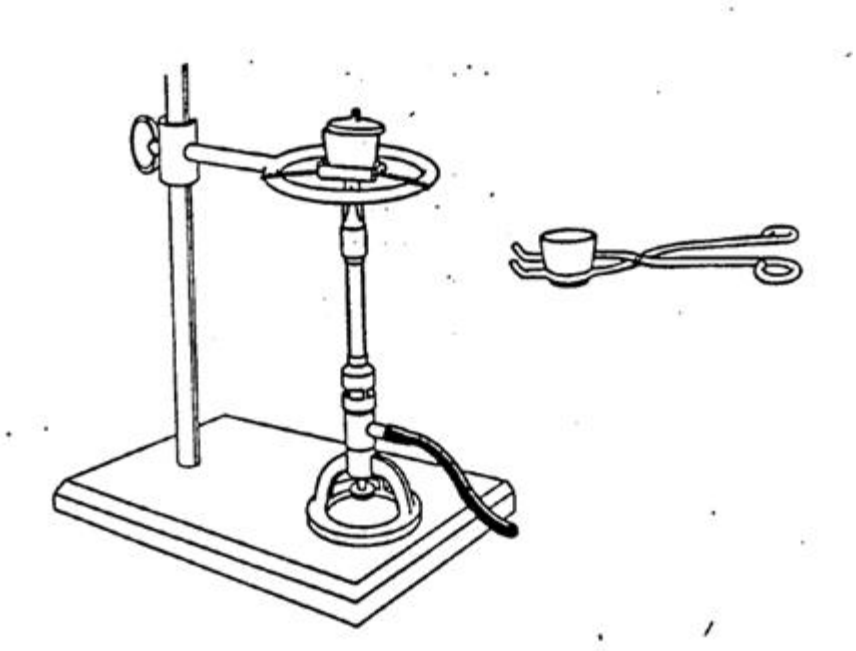
2. Place this in an evaporating dish and place the dish on a clay triangle above a Bunsen burner. Gently heat the solid until the hydrate has been converted to the anhydrous salt. Note what happens during the heating process. Observe the product.

Quantitative Measurments

Now we will make the needed mass measurements to determine the number of Water of hydration in Epsom Salt.

1. Obtain a crucible and lid and wipe it as clean as possible with a piece of paper towel. (Do no other cleaning of the crucible.)

2. Support the crucible/lid on a clay triangle set on a ring above a Bunsen burner.



3. Heat the crucible with an intense flame for about 5 minutes. This will drive any absorbed Water and other combustible materials from the crucible. Allow it to cool for about 10 minutes. (Always handle the crucible and lid with a pair of tongs.)

4. Weigh the empty crucible and lid. This should be done fairly accurately on an analytical balance.

5. Add ~2g of the Epsom Salt to the crucible and re-weigh the crucible and lid.

6. Return the crucible with the sample to the clay triangle and set the lid off the crucible's edge to allow the gaseous Water to escape.

7. At first, heat the sample slowly and then gradually intensify the heat. Do not allow the crucible to become red hot as the anhydrous salt can decompose. Heat the sample for 15 minutes. Cover the crucible with the lid and allow it to cool for about 10 minutes.

8. Re-weigh the sample and the crucible.

Data Analysis

1. Determine the Molar Masses of Water (H_2O) and Magnesium Sulfate ($MgSO_4$).
2. Determine the number Moles Water in your Sample.
3. Determine the number Moles Magnesium Sulfate in your Sample.
4. Determine the number of Hydrating Water molecules *x* per formula unit of Magnesium Sulfate.
5. Report the chemical formula of Epsom Salt.