STOICHIOMETRY AND EMPIRICAL FORMULA

Background

Stoichiometric calculations (calculations of mass relations) are essential for analytical chemistry and/or chemical synthesis in all applications of chemistry. The efficiency of a chemical reaction can be estimated based on its percent yield [% yield = (actual yield/ theoretical yield) x 100]. From results of elemental analysis, percent composition of compounds can be obtained which further can be used to find the simplest or empirical formula of an unknown substance. The empirical formula of a compound gives the simplest whole-number ratio of moles to elements in the compound. Depending on the nature of the compound (ionic or covalent), the empirical formula may or may not be identical with the molecular formula. As ionic compounds do not actually consist of real molecules (we consider the so called formula units, the smallest repeating unit in the crystal lettuce), for these compounds the empirical formula is the usual way to represent them in the chemical formula. Many ionic compounds crystallize with one or more molecules of water loosely bound into the crystalline structure of the anhydrous compound. If the hydrate is colored, a color change usually accompanies the loss of this water of crystallization (or water of hydration) upon heating (efflorescence). Molecular or covalent compounds, on the other hand, do not typically have identical molecular or empirical formulas. To determine their true molecular composition, we need an additional piece of information about the compound: its molar mass (usually obtained by another method , e.g., by mass spectrometry). The ratio of the molar mass to the empirical formula mass provides the multiple factor. Multiplying the subscripts of atoms in the empirical formula by this factor yields the molecular formula.

Purpose

To determine the empirical formula for a hydrate of inorganic ionic compound and it percentage of water by its mass.

Use the following equations:

$$(X) = \text{Moles of anhydrous salt} = \text{Mass of anhydr. salt (g)} \times \frac{\text{1 mol of anhydr. salt}}{\text{Molar mass of anhydr. salt (g)}}$$

$$(Y) = \text{Moles of water} = \text{Mass of water (g)} \times \frac{\text{1 mol of water}}{\text{Molar mass of water (g)}}$$

Divide both, (X) and (Y) by the smallest of the two and find the simplest integer ratio (X) : (Y).

$$\text{Percent by mass of water in the salt} = \frac{\text{Mass of water (g)}}{\text{Mass of (Water + anhydrous salt)(g)}} \times 100$$

Chemicals

A: Nickel (II) sulfate hydrate, (X) $NiSO_4 \cdot (Y)\ H_2O$

B: Iron (III) chloride hydrate, (X) $FeCl_3 \cdot (Y)\ H_2O$

Glassware and Apparatus

50-mL Erlrnmeyer flask (2)

Hot plate
Crucible tongs
Wire gauze
Electronic Balance
Weighing paper or plastic weighing boat
Spatula

Procedure

1. Record the letter of your unknown hydrated salt (A or B).
2. Label 2 clean 50-mL Erlenmeyer flasks for trial 1 and trial 2 and record their masses. Next, record all digits shown on the electronic balance.
3. Add approximately (but exactly!) 1.00 gram of your assigned hydrate to each flask and record the masses.
4. Record the color and appearance of your hydrate on the data sheet.
5. Using the maximum setting on a hot plate, heat both flasks, **Fig. 3.1**, for at least 15 minutes. All condensed water from the mouth of the flasks should evaporate. If needed, heat longer. Heat for additional 5 minutes after all of the water has evaporated. Do not touch the hot plate! ⚠ Keep a safe distance from the hot plate as the liquid may splatter! ⚠

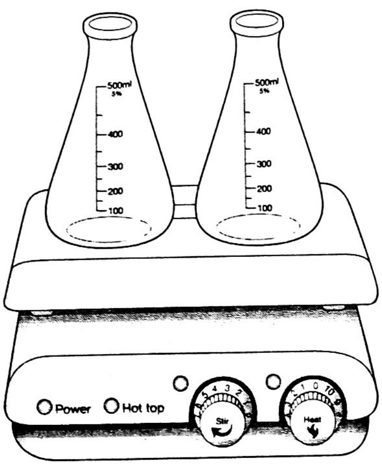

Figure 3.1: Two Erlenmeyer flasks on the hot plate.

6. On your data sheet record both the color and the appearance of the anhydrous salt.
7. Using tongs, remove the flasks from the hot plate and let them cool for approximately 5 minutes on the wire gauze pad on your bench.
8. Weigh the flasks with the anhydrous salt. Record this on your data sheet by subtracting the mass of the empty flask from the mass of the flask with the salt . This will determine the mass of the anhydrous salt (A: $NiSO_4$, or B: $FeCl_3$).
9. Using the Periodic Table of Elements, calculate the number of moles of $NiSO_4$ or $FeCl_3$ and the water .
10. Determine the empirical formula of the hydrate and its water % by using mass.

Disposal: Dispose the anhydrous salts in the appropriate waste containers.

Data and Calculations

Data sheet

	Trial 1	Trial 2
Unknown salt hydrate (A or B)	....	
Color and appearance of unknown hydrate	....	
Mass of unknown hydrate + flask	.... g	.... g
Mass of empty flask	.... g	.... g
Mass of hydrate	.... g	.... g
Color and appearance of unhydrous salt	....	
Mass of unhydrous salt + flask	.... g	.... g
Mass of unhydrous salt	.... g	.... g
Mass of effloresced water	.... g	.... g
Moles of unhydrous salt	.... mol	.... mol
Average moles of unhydrous salt	.... mol	
Moles of water of crystallization	.... mol	.... mol
Average moles of water of crystallization	.... mol	
Simplest integer ratio (X) : (Y)	....	
Empirical formula and name of the original hydrate	....	
Percent by mass of water in the original hydrate	....	

Calculations

<u>Show your step-by-step work in all calculations!</u>

(X) = Moles of anhydrous salt =

Average moles of unhydrous salt =

(Y) = Moles of water of crystallization =

Average moles of water of crystallization =

Divide both, (X) and (Y) by the smaller of the two and find the simplest integer ratio (X) : (Y).

(X) : (Y) =

Percent by mass of water in the original hydrate =