THE MOLE IN CHEMICAL FORMULAS

Lab Overview

Each chemical has its own formula; for example, water is H_2O; hydrogen peroxide is H_2O_2; hydrochloric acid is HCl; ammonia is NH_3. Molecular formulas give the kind and number of atoms (indicated by the subscripts) of each element present in the molecular compound. The **empirical** formula of a compound is the simplest **whole number ratio** of the elements of that compound. In many cases, the molecular formula is the same as the empirical formula. To determine the formula of a substance, we must determine how many grams of each element there are in a sample (of known weight) of that substance. By converting **grams to moles** for each element, we can find a mole-to-mole ratio, which yields the formula.

In this week's lab, you will learn how to derive the chemical formulas of several chemicals by partially decomposing semi-stable substances (called hydrates) to find out how much water they contain. You will then convert the grams of water to moles of water and will be able to calculate a mole-to-mole ratio that will allow you to ultimately determine the formula of the hydrate. You will also observe the formation of a metal oxide and calculate its empirical formula, given appropriate data.

HYDRATES: Some ionic compounds have water (H_2O) incorporated within their formula unit. These compounds are called hydrates and have a specific number of water units within each formula unit of the compound. Note that hydrates are solids, not liquids or solutions, despite containing water. To write the chemical formula of a hydrate, we typically add the specific number of water units per formula unit of compound; this number is just added after the chemical formula of the compound. For example, the hydrate of copper(II) sulfate ($CuSO_4$) has **five water** units associated with each formula unit and it is written as $CuSO_4 \cdot 5H_2O$. The correct name of this hydrate is copper(II) sulfate pentahydrate; the *penta* prefix indicates the presence of the five water units per formula unit of copper(II) sulfate.

Hydrates are not thermally stable and can be dehydrated by heating; the dehydration is a partial decomposition process, in which the water units are released. The class of compounds obtained by the dehydration of hydrates is called *anhydrates*.

The chemical equation of the dehydration of copper(II) sulfate pentahydrate is:

$$\underset{\textit{blue crystals}}{CuSO_4 \cdot 5H_2O} \longrightarrow \underset{\textit{white powder}}{CuSO_4} + 5\,H_2O$$

Another example of hydrates is the chemical compound known as *potassium alum*; the chemical formula of this alum is $KAl(SO_4)_2 \cdot 12\,H_2O$ (name: potassium aluminum sulfate dodecahydrate). This alum is used in water treatment facilities to help gather small particles together for easier removal.

In this week's experiment, you will decompose the alum and one other (unknown) hydrate to find out how much water is in each of them, and thus the mole ratio of the ionic compound to water. You will start by weighing the hydrate before heating. After heating (to drive off the water contained in the hydrate), the resulting anhydrate is weighed to determine how much water has been lost.

Data Processing Examples:

Example 1. A student weighs an empty beaker and finds it to be 23.926 g. She then adds a sample of copper(II) sulfate pentahydrate, and re-weighs it. The combined mass of the beaker and copper(II) sulfate pentahydrate is 25.956 g. After heating, the mass of the beaker and anhydrate is 25.224 g. Determine the mole ratio of water to copper(II) sulfate.

Before Heating	
Empty beaker	$m_1 = 23.926$ g
Beaker + $CuSO_4{\bullet}5H_2O$	$m_2 = 25.956$ g
Mass of $CuSO_4{\bullet}5H_2O$, calculated as $m_3 = m_2 - m_1$	$m_3 = 2.030$ g
After Heating	
beaker + anhydrate	$m_4 = 25.224$ g
Mass of anhydrate ($CuSO_4$), calculated as $m_5 = m_4 - m_1$	$m_5 = 1.298$ g
Moles of $CuSO_4$ (# moles=grams/molar mass of *159.62 g/mol*)	**0.008 moles**
Amount water (H_2O) lost by hydrate (m_3-m_5)	0.732 g
Moles of H_2O (# moles=grams/molar mass of *18 g/mol*) (*This is the amount of water originally in the hydrate sample*)	**0.04 moles**
Ratio of anhydrate to water: $(CuSO_4)_{0.008}\,(H_2O)_{0.04}$ *Since the subscripts are not whole number, we divide both subscripts by the smallest subscript (0.008) to obtain:* $(CuSO_4)_{\frac{0.008}{0.008}}(H_2O)_{\frac{0.04}{0.008}}$ or $(CuSO_4)_1(H_2O)_5$ *Because this compound is a hydrate, the correct formula is written as* $CuSO_4{\bullet}5H_2O$.	

Example 2. A sample of 2.100 g of phosphorus is combined with chlorine, forming a compound with a mass of 9.312 g. Determine the empirical formula of the compound.

*The compound contains both phosphorus and chlorine with a formula of P_xCl_y. To determine its exact formula, we must find **x** and **y**. The compound has a mass of 9.312 g, of which 2.100 g are phosphorus; subtracting phosphorus' mass, we get the mass of chlorine that combined with phosphorus:*	
Mass compound	$m_1 = 9.312$ g
Mass of phosphorus, P	$m_2 = 2.100$ g
Mass of chorine, Cl	$m_2 - m_1 = 7.412$ g
Moles of P (# moles=grams P/atomic weight of P)	**0.07 moles**
Moles of Cl (# moles=grams Cl/atomic weight of Cl)	**0.21 moles**
Ratio of P to Cl: $P_{0.07}\,Cl_{0.21}$ *Since the subscripts are not whole number, we divide both subscripts by the smallest subscript (0.07) to obtain:* $(P)_{\frac{0.07}{0.07}}(Cl)_{\frac{0.21}{0.07}}$ or P_1Cl_3 *Because we don't need to write down a subscript of 1, the correct formula is written as* PCl_3.	

Experimental Procedure

Part A. TA Demonstration Performed in Hood

1. TA places a coil of magnesium in a crucible and heats it to red heat to ignite the magnesium. Then, TA places the lid on the crucible to prevent loss of the product. TA continues to heat until all of the magnesium has been converted to oxide. Students use the data provided in the ***Data Set 1*** to calculate the empirical formula of the magnesium oxide. Have the students start part B of the procedure while working on the calculations.

Part B. Empirical Formula of Alum

1. Clean and dry a 150-mL beaker. Heat it for two minutes. Allow it to cool to room temperature, and then weigh it to ± 0.001 g on the balance. Record this measurement.

2. Add between 1 and 1.5 g of alum to the beaker. Spread the alum to evenly cover the bottom of the beaker. Record the mass of the beaker + alum (Trial 1).

3. Heat the alum gently at first to avoid excessive spattering as the moisture is released. You should observe condensation on the sides of the beaker. You can increase the heating when condensation is visible, but *do not heat* the beaker too intensely, as the alum may decompose too much. When the alum is dehydrated, you should see no remaining moisture on the beaker, and the alum itself will no longer look like individual crystals. Allow the beaker to cool for about 10 minutes.

4. Weigh the beaker with the anhydrate, and calculate the formula of the alum, as shown in ***Example 1***. If your results are satisfactory, dispose of the alum anhydrate in the waste container provided. Repeat this procedure and record your results on the report sheet (Trial 2).

Part C. Empirical Formula of an Unknown Hydrate

1. Clean and dry a 150-mL beaker. Heat it for two minutes and allow it to cool to room temperature. Weigh the beaker to ± 0.001 g. Record the weight of the beaker.

2. Add between 1 and 1.5 g of one of the unknown hydrates to the beaker being careful to spread the unknown evenly over the bottom of the beaker, and record the mass of the beaker and hydrate to ±0.001 g.

3. Heat the hydrate gently at first to avoid excessive spattering as the moisture is released. You should see condensation on the sides of the beaker. You can increase the heating when condensation is visible, but *do not heat* the beaker too intensely, as the alum may decompose too much. When the hydrate is dehydrated, you should see no remaining moisture on the beaker, and the hydrate itself should look like a powder, rather than like crystals. Allow the beaker to cool for about 10 minutes.

5. Weigh the beaker and anhydrate, and calculate the formula of the hydrate, as shown in ***Example 1***. If your results are satisfactory, dispose of the anhydrate in the waste container provided. Repeat this procedure and record your results on the report sheet (Trial 2).

Report Sheet

Name:______________________________Section:__________Date:________________

Data Set 1. Empirical Formula of Magnesium Oxide

Mass of crucible plus magnesium	m_1 = 25.748 g
Mass of crucible	m_2 = 25.503 g
Mass of magnesium, calculated as $m_3 = m_1 - m_2$	m_3 = _______g
Mass of crucible plus magnesium oxide	m_4 = 25.910 g
Mass of reacted oxygen, calculated as $m_5 = m_4 - m_2 - m_3$	m_5 = _______g
Moles of magnesium, Mg (# moles=grams Mg/atomic weight of Mg)	_______ **moles**
Moles of oxygen, O (# moles=grams O/ atomic weight of *O*)	_______ **moles**
Ratio of **Mg** to **O**: Mg_____ O_____ Empirical Formula: Mg O	

Data Set 2. Empirical Formula of the Potassium Alum

	Trial 1	Trial 2
Mass of beaker plus alum, $KAl(SO_4)_2 \cdot$ ***?*** H_2O	g	g
Mass of beaker		
Mass of alum		
Mass of beaker + anhydrate		
Mass of anhydrate, $KAl(SO_4)_2$		
Mass of water		
Moles anhydrate *(The molar mass of $KAl(SO_4)_2$ is 258.21 g/mol)*	_____**moles**	_____**moles**
Moles of water *(The molar mass of H_2O is 18 g/mol)*	_____**moles**	_____**moles**
Trial 1: $KAl(SO_4)_2 \cdot$________H_2O	Trial 2: $KAl(SO_4)_2 \cdot$________H_2O	

Data Set 3. Empirical Formula of an Unknown Hydrate

Formula of the unknown is______________The molar mass of unknown is _______g/mol

	Trial 1	Trial 2
Mass of beaker plus unknown	g	g
Mass of beaker		
Mass of unknown		
Mass of beaker + anhydrate		
Mass of anhydrate		
Mass of water		
Moles anhydrate *(Molar mass of unknown __________g/mol)*	_____**moles**	_____**moles**
Moles of water *(Molar mass of H_2O is 18 g/mol)*	_____**moles**	_____**moles**
Trial 1: __________·________H_2O	Trial 2: _________·________H_2O	

POST-LAB ASSIGNMENT #4: THE MOLE IN CHEMICAL FORMULAS

Solve the problem below and upload your solution to bblearn

Date:______________________

Name:_____________________

Section:____________________

A sample of copper (Cu) wire with a mass of 1.659 g is heated with excess powdered sulfur (S) until the copper has combined with the sulfur, and all the extra sulfur has been burned off. The final product (a sulfide of copper, Cu_xS_y) has a mass of 2.080 g. What is the empirical formula of the product?