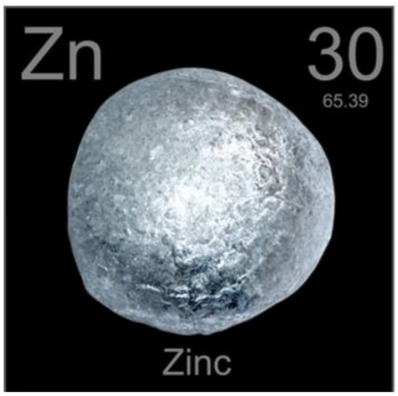
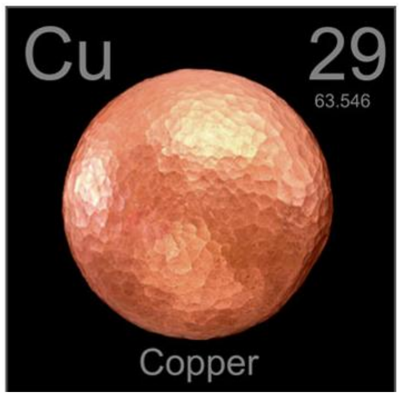
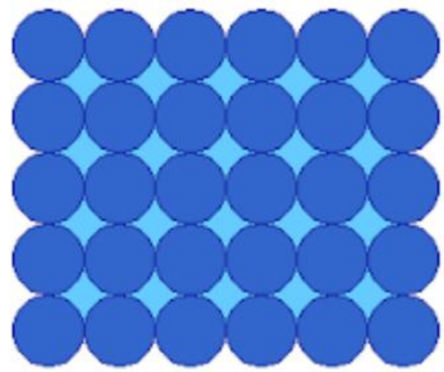
CEAC 105
GENERAL CHEMISTRY

Experiment 4

Determination of The Atomic Weight of a Metal







Purpose: To determine atomic weight of a metal by measuring the volume of the hydrogen gas liberated from the oxidation of a metal with an acid reaction.

MATERIALS:

A piece of metal	Beaker	Methyl orange
Cotton	Burette	Hydrochloric acid (HCl)

THEORY:

A metal, which is a good reducing agent, will react with hydrochloric acid, liberating hydrogen gas and forming a salt.

M (solid) + nH^+ (aqueous) --- > M^{n+} (aqueous) + $n/2\ H_2$ (gas)

In this experiment, a weighted sample of metal will be completely dissolved in HCI solution, and the volume of hydrogen liberated will be determined from the volume of water that it displaces. From this data the atomic weight of the metal will be calculated. This calculated value is the experimental value of the atomic weight of the metal.

After the determination of the atomic weight of the metal, your assistant will tell you the name and the atomic weight of the metal (theoretical value). Now, you can calculate percentage error of atomic weight calculation. The valence of the metal is taken as n+ for the reaction given above.

A mole is defined as the amount of substance that contains the same number of chemical units as the number of atoms in exactly 12.0000 g of ${}^{12}_{6}C$

1 mole ${}^{12}_{6}C$ atoms = 12.0000 g ${}^{12}_{6}C$

A mole of atoms, therefore, consists of Avogadro's number of atoms and has a mass in grams numerically equal to the atomic weight of the element, shortly.

$$\text{Number of moles of an atom} = \frac{\text{weight}}{\text{Atomic Weight}}$$

$$\text{Number of moles of a molecule} = \frac{\text{weight}}{\text{Molecular Weight}}$$

Before doing this experiment, it will be necessary for the student to review his knowledge about the ideal gas equation.

Ideal Gas Equation:

The behavior of an ideal gas is expressed by the well-known equation given below.

PV = nRT   Ideal Gas Equation

In the equation the units are under STP (1 atm and 0 °C, 273 K) 1 mole of ideal gas occupies 22.4 L volume. So;

$$R = \frac{(1\ atm) \times (22.4\ L)}{(1\ mol) \times (273\ K)} = 0.00823\ {}^{L \times atm}\!/_{mol \times K}$$

EXAMPLE 1:

Standard conditions are defined as 0^oC and 1 atm for temperature and pressure, respectively. Using ideal gas equation, calculate the volume that is occupied by one mole of an ideal gas under the standard conditions (STP).

SOLUTION:

$T = 0^oC$

$T = 0^oC + 273.15 = 273.15$ K

PV = nRT V = nRT/P

V= (1 mole)*(0.082 L.atm.mole^{-1}K^{-1})*(273 K) / 1 atm

V=22.4 L/mole

EXAMPLE 2:

In a reaction, 334 mL of H_2 gas is liberated at a pressure of 674 mm Hg and a temperature of 17^oC. Calculate the number of moles of H_2 gas liberated.

SOLUTION:

760 mm Hg = 1 atm

P = (674 mm Hg)*(1 atm /760mm Hg) = 0.887 atm

$T = 17^oC + 273.15 = 290.15$ K

V = 0.334 L

R = 0.082 L.atm.mole^{-1}K^{-1}

$$n = \frac{PV}{RT} = \frac{0.887 \text{ x } 0.334}{0.082 \text{ x } 290.15} = 0.012 \text{ moles of } H_2$$

EXAMPLE 3:

0.26 g of an unknown metal is liberating 0.004 moles of H_2 gas. If the valence of this metal is 2+, find the atomic weight of this metal. Find the name of this metal using the periodic table.

X(solid) + nH^+ (aqueous) --- > X^{n+} (aqueous) + n/2 H_2 (gas)

SOLUTION:

According to the stoichiometry of the reaction;

Number of moles of metal equals to the number of moles of the metal

$$\text{Number of moles} = \frac{\text{weight}}{\text{Atomic Weight}} \longrightarrow \quad \text{At. Wt.} = \frac{\text{weight}}{\text{number of moles}} = \frac{0.26}{0.004} = 65.0 \text{ g/mole}$$

Closest value to this experimental atomic weight value is 65.39, thus the metal is predicted to be Zinc.

Table 1: Conversion of pressure units

Units	Pa	psi	atm	bar	torr
Pa	1N/m^2	1.45 x 10^{-4}	9.869 x 10^{-6}	10^{-5}	7.5 x 10^{-3}
psi	6.894 x 10^3	1 lb/in^2	6.8 x 10^{-2}	6.894 x 10^{-2}	51.714
atm	1.01325 x 10^5	14.695	P_0	1.01325	760
bar	10^5	14.5	0.9869	10^6 dyne/cm^2	750
torr	133.322	1.93 x 10^{-2}	1.315 x 10^{-3}	1.333 x 10^{-3}	1 mmHg

Dalton's Law of Partial Pressure

Dalton's Law (also called Dalton's Law of Partial Pressures) states that the total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of individual gases. Mathematically, this can be stated as follows:

$$P_{total} = P_1 + P_2 + P_3 + \cdots + P_n$$

where P_1, P_2 and P_n represent the partial pressures of each compound. It is assumed that the gases do not react with each other.

Collecting Gases Over Water

The amount of gas present can be determined by collecting a gas over water and applying Dalton's Law.

EXAMPLE 4:

O_2 gas is collected in a pneumatic trough with a volume of 0.155 L until the height of the water inside the trough is equal to the height of the water outside the trough. The atmospheric pressure is 754 torr, and the temperature is 295 K. How many moles of oxygen are present in the trough? (At 295 K, the vapor pressure of water is 19.8 torr.)

SOLUTION:

The total pressure in the tube can be written using Dalton's Law of Partial Pressures:

$$P_{total} = P_{H_2O} + P_{O_2}$$

Rearranging this in terms of P_{O_2}, we have:

$$\mathrm{P}_{\mathrm{O}_2} = P_{total} - P_{H_2O}$$

Because the height of the water inside the tube is equal to the pressure of the water outside the tube, the total pressure inside and outside the tube must be equal to the atmospheric pressure. With substitution, we have:

$$\mathrm{P}_{\mathrm{O}_2}=\mathrm{P}_{\mathrm{total}}-\mathrm{P}_{\mathrm{H_2O}}=754-19.8=734 \text{ torr}=0.966 \text{ atm}$$

Next, we apply the Ideal Gas Law:

$$n=\frac{P\times V}{R\times T}=\frac{(0.966\text{ atm})\times(0.155L)}{(0.082\ {}^{L\times atm}/_{mol\times K})\times(295K)}=0.00619\text{ mol } O_2$$

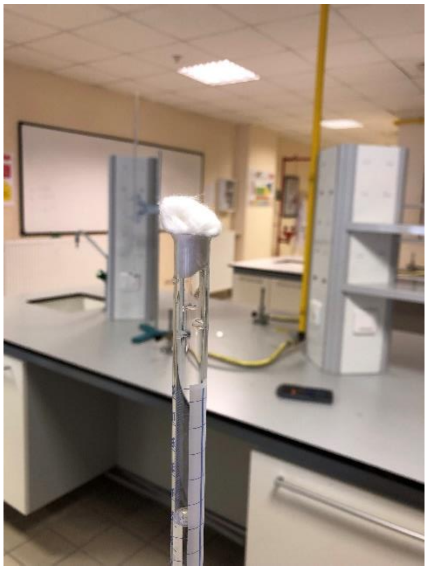
Table 2: Partial pressure values of water at different temperature values

T(°C)	P(torr)
17	14.5
18	15.5
19	16.5
20	17.5
21	18.7
22	19.8

T(°C)	P(torr)
23	21.1
24	22.4
25	23.8
26	25.2
27	26.7
28	28.4

<table>
<tr><th colspan="2">PROCEDURE:</th></tr>
<tr><td>1. Mix 20 mL of 3M HCl and 1 drop of methyl orange and pour it into a burette.</td><td></td></tr>
<tr><td>2. Add distilled water up to 5 mL scale of same burette.</td><td></td></tr>
</table>

3. Weigh out about 1 cm length of your metal. Record this weight.	
4. Take the sample and place it in a small bag made from a piece of cotton. (Wrap around the metal by a piece of cotton.)	

5. Tie this metal in cotton with a piece of thread and hang it a few centimeters above from the top of the liquid in the burette, providing that it does not come into contact with the acid solution. You may drop it into the solution as well.	
6. Make sure your set-up look like in the figure.	

7. Put some water in the beaker and set up the apparatus as shown in the figure, read and record the first height of pink acidic solution (V_1).	
8. Make sure the cotton is stacked at the end of burette. When the system is shown to be airtight, wait until the acid solution comes in contact with the metal. You will see the hydrogen gas is being produced as bubbles near the metal, and some water will pour into the beaker.	

9. After the metal has completely dissolved, read the volume of the water in the burette and record it (V_2). The displaced amount of water in mL will be taken as the amount of hydrogen gas liberated. Record the room temperature. Record the atmospheric pressure, P_{atm} and the vapor pressure of water, P_{H_2O} at this temperature.	

10. Calculate the partial pressure of H_2 gas, P_{H_2} , in the flask; express it in the units of atm.

11. Calculate the number of moles of H_2.

12. What is the number of moles of metal?

13. Calculate the atomic weight of metal.

14. Ask your **assistant the name of the metal.**

15. Ask your assistant the exact atomic weight of the metal.

16. Calculate the percent error for the experimental atomic weight of the metal.

DATA SHEET

Determination of The Atomic Weight of a Metal

Student's Name:

Laboratory Section/Group Number:

Assistant's name and signature: Date:

DATA

1. Weight of the metal (W) =…………………g

2. Initial height of the water (V_1) =……………… mL

3. Final height of the water (V_2) =……………… mL

4. Volume of the water displaced ($V_{H_2O} = V_1 - V_2$) =…………………mL

5. Volume of H_2 gas produced ($V_{H_2} = V_{H_2O}$) =…………………L

6. Temperature ($T_{room} = T_{H_2}$) =…………………^oC

7. Vapor pressure of H_2O at T = P_{H_2O} (Use Table 2) =…………………mm Hg

8. Atmospheric pressure= P_{atm} (ask your assistant) =…………………mm Hg

9. Partial pressure of H_2 gas $P_{H_2} = P_{atm} - P_{H_2O}$ =…………………mm Hg

10. Number of moles of H_2, $n_{H_2} = \frac{P_{H_2}V_{H_2}}{RT_{H_2}}$ =…………………

11. Number of moles of metal =…………………

12. Atomic Wt $= \frac{\text{Weight}}{\textit{number of moles of metal}}$ = ………………… g/mole

13. Name of the metal =…………………

14. Percent error for experimental Atomic Weight =…………………%

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Experiment 5
Stoichiometry: The Reaction of Iron with Copper(II) Sulfate

Purpose: To enhance the understanding of stoichiometry, a reaction between iron and copper (II) sulfate ($CuSO_4$) solution will be conducted. This will help you to differentiate limiting and excess reactant in a chemical reaction. Finally the theoretical and percent yield of this reaction will be calculated.

Theory

Stoichiometry is the measurement of quantitative relationships in chemical formulas and equations.

Mostly in chemical reaction two or more reactants (reagents) are placed. The reagent that is totally consumed before other reactants at that time reaction stops defined as **limiting reagent**. Limiting reagent is the reactant which is totally consumed before other reactants. The limiting reactant limits the amount of product that can be formed since the consumption of the reactant stops the reaction.The reactant that is remained once the limiting reagent is consumed is called as **excess reagent**.

Example : How many moles of Ag are in 1.75 mol of Ag_2CrO_4 ?

There are 2 moles of Ag atoms for each Ag_2CrO_4 formula unit

Therefore,

1.75 mol Ag_2CrO_4 x 2 = 3.5 mol Ag

Example: Sodium carbonate**,** Na_2CO_3, reacts with calcium chloride, $CaCl_2$, to form calcium carbonate, $CaCO_3$ and sodium chloride, NaCl . In an experiment 53 grams of Na_2CO_3 and 44.4 grams of $CaCl_2$ were mixed and 23.6 gr $CaCO_3$ was obtained. Which is the limiting reactant? What is the percent yield of $CaCO_3$?

(Molecular weight of Na_2CO_3 = 106 g/mole, $CaCl_2$ =111 g/mole, $CaCO_3$ = 100 g/mole, NaCl=58.5 g/mole)

$$Na_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2NaCl(aq)$$

Mole of Na_2CO_3 : 53/ 106 = 0.3 moles of Na_2CO_3 are present.

Mole of $CaCl_2$: 44.4/ 111 = 0.4 moles of $CaCl_2$ are present.

1 mole of Na_2CO_3 reacts with 1 mole of $CaCl_2$ according to the above equation.

$$\frac{1 \text{ mole } Na_2CO_3}{0.3 \text{ mole } Na_2CO_3} = \frac{1 \text{ mole } CaCl_2}{x}$$

x = 0.3 moles of $CaCl_2$ are needed.

0.3 moles of $CaCl_2$ are needed and 0.4 moles of $CaCl_2$ are present. Therefore, $CaCl_2$ is excess reagent and Na_2CO_3 is limitng reactant.

Since Na_2CO_3 is limiting reactant it limits the amount of the product, $CaCO_3$, that will be formed.

$$\frac{1 \text{ mole } Na_2CO_3}{0.3 \text{ mole } Na_2CO_3} = \frac{1 \text{ mole } CaCO_3}{x}$$

x = 0.3 moles of $CaCO_3$ will be formed

0.3* 100= 30 gr of $CaCO_3$ will be formed. This amount is theoretical yield. Since the actual yield is the amount of product given, the percent yield will be calculated by the formula given below;

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Then, percent yield for $CaCO_3$ is;

(23.6/30)*100 =78.7%

In this experiment stoichiometric principles will be used to obtain the appropriate equation between the reaction of iron metal and copper(II) sulfate solution. Whent the reaction starts, the formation of metallic copper, which is precipitating during reaction as a finely divided reddish-orange powder will be observed. This reaction is one of the example of single substitution reaction in which one element “displaces” with another element in a compound The element which has ability of displacing other element from a compound is said to be “more active” than the displaced one. In this experiment, iron is more active than copper.

Two distinct forms of iron are present, namely Ferrous, $Fe^{2+}$and Ferric, Fe^{3+}. Stoichiometric principles will be used to determine which reaction is more dominant compared to other by examining the reaction between iron and copper (II) sulfate solution. If Fe^{2+} is formed, then **equation (1)** is dominant, but **equation (2)** will be selected if Fe^{3+} is formed. This can be determined by taking the mole ratio of copper to iron. If the moles of copper is equal to the moles of iron, then equation (1) has taken place. If you obtain 1.5 moles of copper per mole of iron, in this case equation (2) should be selected Find out which equation is corresponding to the results of the experiment you have done.

Equation (1) $Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$, other representation is

$$Fe_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Fe^{2+}{}_{(aq)} + Cu_{(s)}$$

Equation (2) $2\,Fe_{(s)} + 3\,CuSO_{4\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(aq)} + 3\,Cu_{(s)}$,

other representation is

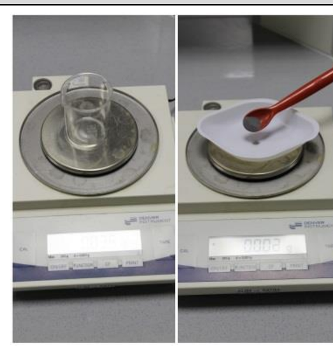
$$2\,Fe_{(s)} + 3\,Cu^{2+}{}_{(aq)} \rightarrow 2\,Fe^{3+}{}_{(aq)} + 3\,Cu_{(s)}$$

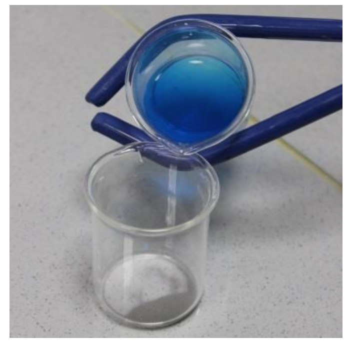
To the known amount of iron, excess of copper (II) sulfate solution will be added. The purpose of using excess solution is owing to provide the complete reaction of iron. The metallic copper produced will be weighed after washing and drying processes and these weighings will be used to calculate the moles of iron used and the moles of copper formed at the end of the raection.

Materials

Fe powder	Acetone	Glass stick
Cupper (II) Sulfate ($CuSO_4$)	Beaker	Bunsen burner

Procedure

The Reaction of Iron with Copper(II) Sulfate	
1. Weigh a dry and clean 100 or 250 mL beaker and record the weight of it onto your data sheet. then, accurately weigh 1.00 gram of iron powder into this beaker. Do not exceed 1.01 grams.	

2. Measure 30 mL of 1.0 M $CuSO_4$ solution by using a graduated cylinder. Pour this solution into another beaker, and heat gently to almost boiling.	
3. Slowly add hot $CuSO_4$ solution to the beaker that contains the iron powder. Stir the mixture a few times until completeness of the reaction. You should see copper forming. When the reaction has finished, allow the copper product to cool.	
4. Then carefully decant the liquid from the copper into the waste container. Be careful not to lose any copper.	
5. Add about 10 mL of distilled water to the solid copper and swirl to wash any remaining ions from the copper. Decant the wash water from the copper and add 10 more mL of distilled water, swirl and decant again. Wash copper particles finally with several mL of acetone (**CAUTION**-Acetone is very flammable). Swirl and allow to stand a few minutes. Decant off the acetone.	

6. The acetone readily dissolves the water and helps the removal of it from the medium. Swirl the beaker gently on low heat flame. Copper product should be spread in a single layer on the bottom of the beaker. Grinding of aggregates with a spatula makes the copper easy to dry. Be sure not to remove any copper from the beaker.	
7. After drying, allow copper to cool and weigh the beaker plus copper to calculate the mass of copper formed. Record the mass on your data sheet. Finally, calculate the moles of iron used and the moles of copper formed to determine which reaction of iron is taking place, reaction (1) or reaction (2).	

DATA SHEET

Stoichiometry: The Reaction of Iron with Copper(II) Sulfate

Student's Name : Date:

Laboratory Section/Group No :

Assistant's Name and Signature:

Data and Calculations

Mass of empty beaker :

Mass of iron used :

Moles of iron used :

Mass of beaker plus copper :

Mass of copper formed :

Moles of copper formed :

Moles of Cu divided by moles of Fe :

Reaction Equation :

Limiting Reagent :

Theoretical and Percent Yield :

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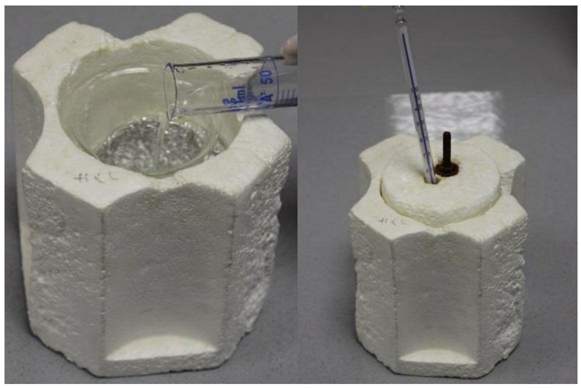
Experiment 6
Heat of Neutralization



Purpose: To calculate enthalpy change of a reaction by using calorimeter and understand the difference between endothermic and exothermic reactions

Prelaboratory Work

Before the experiment in the laboratory, you should be able to answer these questions.

1. Calculate the volume of solution of 1.50 M HNO_3 required in part (A) of the procedure.

2. Calculate the mass of NH_4NO_3 required in part (B) of the procedure. Show your reasoning.

3. Explain in stepwise fashion (number the steps) how you will determine the ΔH of formation of NH_4NO_3 from your data and information given in the experiment. Use actual numbers whenever possible.

4. How many joules are required to change the temperature of 80.0 g of water from 23.3°C to 38.8 $^{\circ}$C?

Theory

In chemical reactions, energy change is observed. This energy change is usually in the form of heat and at constant pressure it is defined as **heat of reaction** or **enthalpy change (ΔH).** To form 1 mole of compound from its constituent elements, necessary amount of enthalpy change occurs and this change is defined as **enthalpy of formation**. If heat is released during the reaction, ΔH is shown with negative sign and the reaction is called **exothermic reaction.** If heat is absorbed during the reaction, ΔH is shown with positive sign and the reaction is called **endothermic reaction.**

Direct measurement of enthalpies of formation is difficult experimentally, so indirect methods involving enthalpies of reaction are used. Hess's Law states that the change in a thermodynamic property such as enthalpy depends on the initial and final states and is independent of path followed. An example for Hess's Law is given below.

$\frac{1}{2} N_2 (g) + 3/2 H_2 \rightarrow NH_3 (g)$ $\Delta H_{formation} (NH_3(g)) = \Delta H_1$

$NH_3 (g) \rightarrow NH_3 (aq)$ $\Delta H_{dissolving} = \Delta H_2$

Assume that ΔH_1 and ΔH_2 are known. If first and second reactions are added, net reaction becomes;

$\frac{1}{2} N_2(g) + 3/2 H_2 \rightarrow NH_3(aq)$ $\Delta H_{formation} (NH_3(aq)) = \Delta H_1 + \Delta H_2$

which is also formation reaction of $NH_3(aq)$.

Given that ΔH_1 = - 45.8 kJ/mol and ΔH_2 = - 35.4 kJ/mol, we can calculate the ΔH formation of $NH_3(aq)$ as -81.2 kJ/mol.

The heat is measured experimentally by allowing the reaction to take place in a thermally insulated vessel called as **calorimeter**. If the calorimeter is perfectly insulated, no heat change occurs between system and surrounding and the system is defined as adiabatic (Q=0). Consequently, at constant pressure, ΔH_{system} is also equal to zero. The formulation of enthalpy change of the system, ΔH_{system}, is shown as in Equation (1).

$$\Delta H_{system} = \Delta T \text{ (heat capacity of calorimeter + heat capacity of contents)} \quad (1)$$

For endothermic reaction in adiabatic system, Equation (1) can be written as below.

$$\Delta H_{system} = n\Delta H_{reaction} + C_p\Delta T$$

$$0 = n\Delta H_{reaction} + C_p\Delta T$$

$$n\Delta H_{reaction} = - C_p\Delta T$$

In a similar manner, for exothermic reaction in an adiabatic system, Equation (1) can be simplified as:

$$\Delta H_{system} = -n\Delta H_{reaction} + C_p\Delta T$$

$$0 = -n\Delta H_{reaction} + C_p\Delta T$$

$$n\Delta H_{reaction} = C_p\Delta T$$

In this experiment, you will determine the heat of formation of various ammonium salts $NH_4X(s)$ where X is Cl, NO_3 or SO_4 by combining measurements of the heat for the neutralization reaction;

$NH_3(aq) + HX(s) \rightarrow NH_4X(aq)$ ΔH_{neut}

And the heat of the dissolution reaction;

$NH_4X(s) + H_2O \rightarrow NH_4X(aq)$ ΔH_{diss}

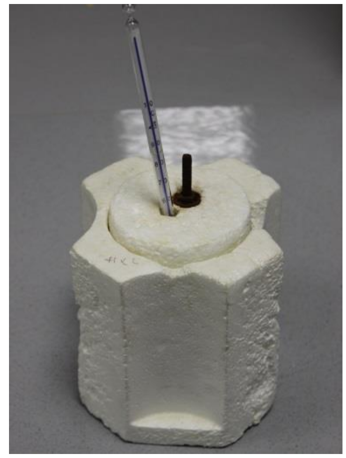
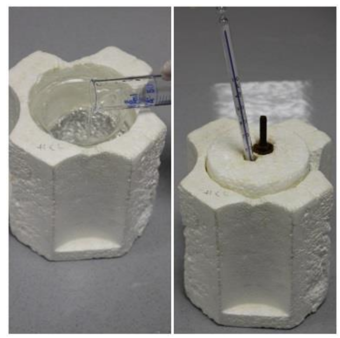
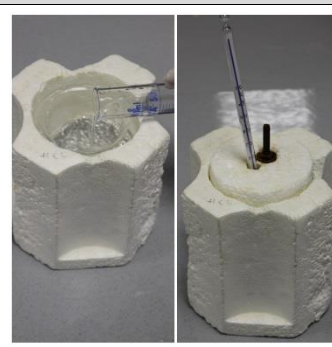
with known heats of formation of $NH_3(aq)$ and $HX(aq)$.

Materials

Nitric acid (HNO_3)	250 mL beaker	Graduated cylinder
Ammonia (NH_3)	Thermometers	
Ammonium nitrate (NH_4NO_3)	Styrofoam cups	

Procedure

PART A: Heat of Neutralization	
1. Obtain a sytrofoam cup. In the first cup, place 50 mL of 1.5 M NH_3.	

2. Place a thermometer in the cup containing the NH_3 and record temperature at 30 seconds intervals.	
3. Add the acid solution to the NH_3 and swirl to mix. Continue taking temperature data at 30 seconds intervals while swirling the solution occasionally.	
PART B: Dissolving	
1. Place a volume of distilled water equal to the final volume of solution from part (A) in a Styrofoam cup and record temperature data at 30 seconds intervals.	
2. Weigh out that mass of NH_4NO_3 salt into a clean, dry beaker.	

3. Immediately, add the weighed amount of salt, swirl to dissolve (use stirring rod if necessary), and continue taking temperature data at 30 seconds intervals.	

Calculations

1. Plot temperature versus time graph using your data and determine ΔH_{neut} for (a) and ΔH_{diss} for (b).

2. Take ΔH_f of 1.5 M NH_3 as -81.2 kJ/mol and ΔH_f of 1.5 M HCl as -165.1 kJ/mol, calculate the ΔH_f of $NH_4Cl(s)$.

3. Do the same calculations for $NH_4NO_3(s)$ using -206.0 kJ/mol for the ΔH_f of 1.5 M HNO_3.

4. Calculate ΔH_f of $(NH_4)_2SO_4$ using -884.2 kJ/mol for the ΔH_f of 1.5 M H_2SO_4. [Note that all ΔH_f are per mol (not per 1.5 mol)]

DATA SHEET

Heat of Neutralization

Student's Name : Date:

Laboratory Section/Group No :

Assistant's Name and Signature :

Prelaboratory Work

1. Volume of 1.5 M HNO_3 solution:

2. Mass of NH_4NO_3:

Time (s)	Temperature of NH_3 solution in °C before adding 1.5 M HNO_3	Temperature of distilled water in °C before NH_4NO_3
0		
30		
60		
90		
120		
150		
180		
Time (s)	Temperature of NH_3 + 1.5 M HNO_3 solution in °C	Temperature of NH_4NO_3 solution in °C
0		
30		
60		
90		
120		

150		
180		

Results

1. Calculate $\Delta H_{neutralization}$:

2. Calculate $\Delta H_{dissociation}$:

3. Calculate $\Delta H_{formation}$ of NH_4NO_3: