CEAC 105
GENERAL CHEMISTRY

Experiment 4
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/mol, $CaCl_2$ =111 g/mol, $CaCO_3$ = 100 g/mol NaCl=58.5 g/mol)

Na_2CO_3(aq) + $CaCl_2$(aq) $CaCO_3$(s) + 2NaCl(aq)
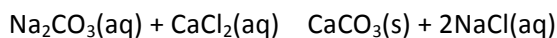
Mole of Na_2CO_3 : 53/ 106 = 0.3 moles of Na_2CO_3 are present.

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 limiting 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. When 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 are 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) $\qquad Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$, other representation is

$$Fe_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Fe^{2+}{}_{(aq)} + Cu_{(s)}$$
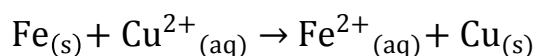
Equation (2) $\qquad 2\,Fe_{(s)} + 3\,CuSO_{4(aq)} \rightarrow Fe_2(SO_4)_{3(aq)} + 3\,Cu_{(s)}$, other representation is

$$2\,Fe_{(s)} + 3\,Cu^{2+}{}_{(aq)} \rightarrow 2\,Fe^{3+}{}_{(aq)} + 3\,Cu_{(s)}$$
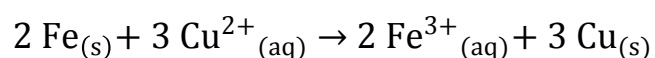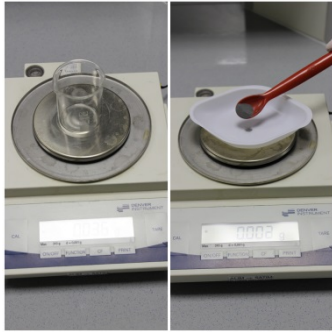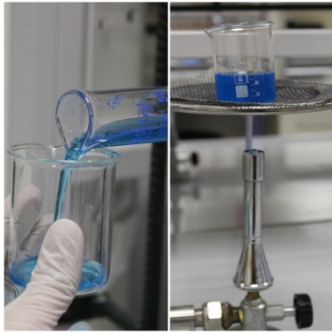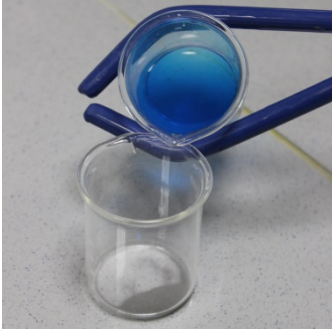
To the known amount of iron, excess of copper (II) sulfate solution will be added. The purpose of using excess solution is owing to provide the complete reaction of iron. The metallic copper produced will be weighed after washing and drying processes and these weightings will be used to calculate the moles of iron used and the moles of copper formed at the end of the reaction.

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 weigth 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 standing 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.	
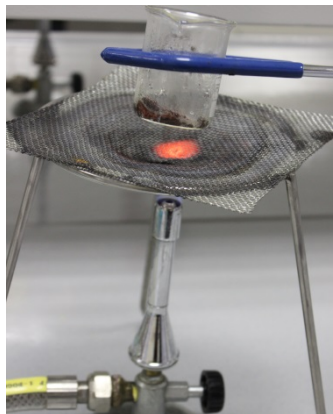
7. After drying, allow copper to cool and weigh the beaker plus copper to calculate the mass of copper formed. Record the mass on your data sheet. Finally, calculate the moles of iron used and the moles of copper formed to determine which reaction of iron is taking place, reaction (1) or reaction (2).

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   :

CEAC 105
GENERAL CHEMISTRY

Experiment 5
Molecular Weight Determination from Freezing Point Depression

Purpose: To learn colligative properties and to determine the molecular weight of sulfur by using one of the colligative properties.

Prelaboratory Work

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

1. List the colligative properties.
2. What is the molality of a solution that contains 1.5 g of urea (molar mass = 60 g/mol) in 200 g of benzene?
3. Calculate the freezing point of a solution containing 5.85 g of NaCl in 200 g of water? (Na: 23 g/mol; Cl:35.5 g/mol)
4. A solution containing 1.00 g of an unknown substance in 12.5 g of naphthalene was found to freeze at 75.4°C. What is the molar mass of the unknown substance?
5. Define molality and molarity.

Theory

Some of the physical properties of solutions are independent of the nature of the solute and depend only upon the solute amount. These properties are called "colligative properties". Vapor pressure lowering, Boiling point elevation, Freezing-point depression, and osmotic pressure are colligative properties.

Those effects introduced above are familiar to us in our daily life. In this experiment, freezing point depression property will be used. That's why it is better to give examples or explanations related with this property.

Anti-freeze solution is used for lowering the freezing point of water in automobiles. (This solution is made by mixing water and ethylene glycol or water and methanol or water and another one.) When something is added to water, freezing point of water decreases automatically. The increasing addition amount decreases the freezing point more.

Another example is pouring salt onto the snow on the roads to prevent freezing in winter days. When salt is added to water, the resulting solution has lower freezing point (and also higher boiling point). Increasing the amount of salt decreases the freezing point more. Freezing point and boiling point of a solution differ from those of the pure solvent by amounts that are directly proportional to the molal concentration of the solute and can be expressed with the equation below:

$$\Delta T = K \cdot m$$

where ΔT is the freezing-point lowering or boiling-point elevation, **K** is molal freezing (or boiling) constant which is specific for each solvent, **m** is the molality of the solution (number of moles of solute /kg of solvent).

Some representative constants, boiling points, and freezing points are given in Table 1. For naphthalene, the solvent used in this experiment, the molal freezing constant (K_f) has a value of 6.90 ^{o}C/m.

TABLE 1. Molal Freezing Point and Boiling Point Constants

Solvent	Freezing point (°C)	K_{fp} (°C/m)	Boiling point (°C)	K_{bp} (°C/m)
CH_3COOH (acetic acid)	16.6	3.90	118.1	2.93
C_6H_6 (benzene)	5.4	5.12	80.2	2.53
C_2H_5OH (ethyl alcohol)	-114.1	-	78.4	1.22
H_2O (water)	0.0	1.86	100.0	0.51
$C_{10}H_8$(naphthalene)	80.6	6.9	218	-

Example: What would be the freezing point of a solution containing 19.5 g of biphenyl ($C_{12}H_{10}$) dissolved in 100 g of naphthalene if the normal freezing point of naphthalene is 80.6°C?

Solution:

$$\text{moles of } C_{12}H_{10} = \frac{19.5 \text{ g}}{154 \text{ g/mol}} = 0.127 \text{ mol}$$

$$\frac{\text{moles of } C_{12}H_{10}}{\text{kg of naphthalene}} = \frac{0.127 \text{ mol}}{0.1 \text{ kg}} = 1.27 \text{ molal (m)}$$

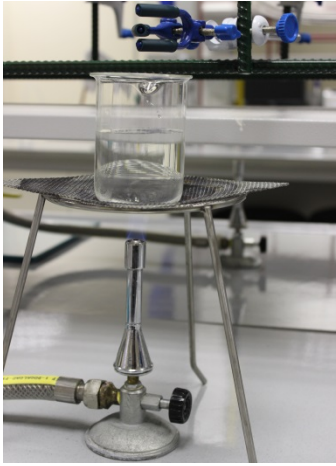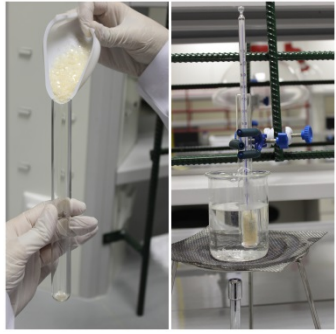$$\Delta T = (6.91 \text{ °C/m}) \times (1.27 \text{ m}) = 8.8 \text{ °C}$$

Since the freezing point is lowered, the observed freezing point of the solution will be $T_{pure} - T_{solution} = 80.6°C - 8.8°C = 71.8°C$

Materials

Naphthalene	Large test tube	600-mL beaker
Sulfur	Clamp	Bunsen burner
Thermometer	Ring and ring stand	

Procedure

PART A: Cooling Curve for Pure Naphthalene	
1. Fill your 600-mL beaker nearly full of water and heat it to about 85 $^{\circ}$C. Clamp the test tube in the water bath as shown in figure.	
2. Weigh 3 g of naphthalene and put the naphthalene in the test tube. Then insert the thermometer into the test tube-naphthalene as shown in figure above. Wait for naphthalene melts. When all of the naphthalene has melted, stop heating; remove the test tube out of water.	

|	
3. Starting with 85 ^{o}C; record temperature readings every 30 s. Shake the test tube slightly at every temperature reading. When the freezing point is reached, crystals will start to form, and the temperature will remain constant. After all the naphthalene freeze, temperature lowering will start again.	
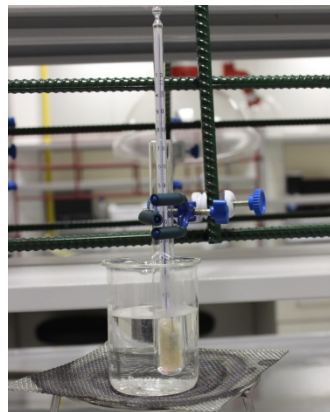
**PART B: Determination of the Molar Mass of Sulfur**

1. Weigh 0.5 g of sulfur.	
2. Replace the test tube in the water bath and heat until all the naphthalene has melted. Add the sulfur into the test tube. Shake the test tube until all the sulfur has dissolved.	
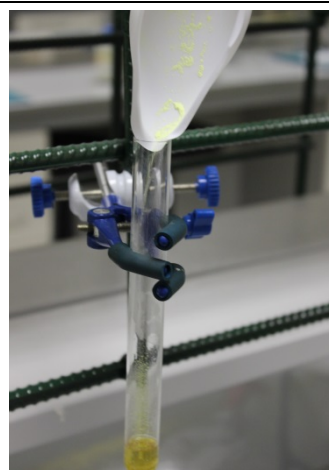
3. Remove the test tube out of the water bath, record the temperature every 30 s until all the naphthalene has solidified. Do not forget to shake the test tube at each temperature readings. Using the data above, draw two cooling graph and find ΔT.
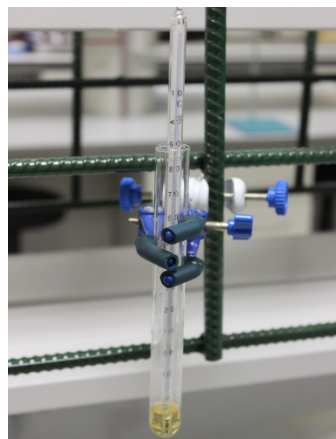
DATA SHEET

Molecular Weight Determination from Freezing Point Depression

Student's Name   :   Date:

Laboratory Section/Group No :

Assistant's Name and Signature:

1. Weight of Naphthalene: …………..g

2. Weight of Sulfur:………………..g

3. Freezing point of pure naphthalene, T_{pure} :……………………… ^{o}C

4. Freezing point of solution (napht. + sulfur), $T_{solution}$:……………. ^{o}C

5. $\Delta T = T_{pure} - T_{solution}$ = ……………..^{o}C

6. Molality of Sulfur: …………………..molal (m). *(Show calculations below..)*

7. Molar mass of unknown Sulfur:…………………..g/mol *(Show calculations below..)*

COOLING CURVE

Pure naphtalene		Naphtalene + Sulfur (S)
Temperature (°C)	Time (sec.)	Temperature (°C)

CEAC 105
GENERAL CHEMISTRY

Experiment 6
Heat of Neutralization
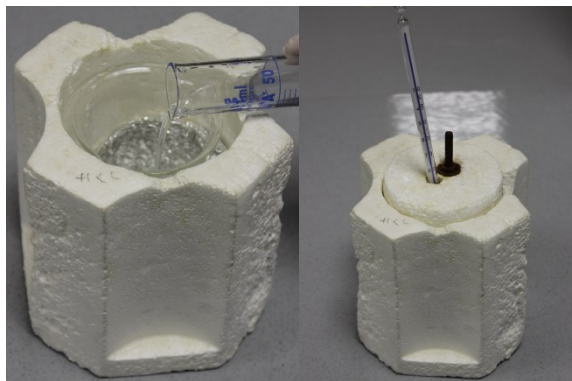
Purpose: To calculate enthalpy change of a reaction by using calorimeter and understand the difference between endothermic and exothermic reactions

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 $38.8 °C$?

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.

½ N_2 (g) + 3/2 H_2 → NH_3 (g) $\Delta H_{formation}$ (NH_3(g))$= \Delta H_1$

NH_3 (g) → 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;

½ N_2(g) + 3/2 H_2 → 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 $\Delta H_1 = $ - 45.8 kJ/mol and $\Delta H_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).

$\Delta H_{system} = \Delta T$ (heat capacity of calorimeter + heat capacity of contents)   (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}
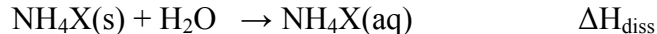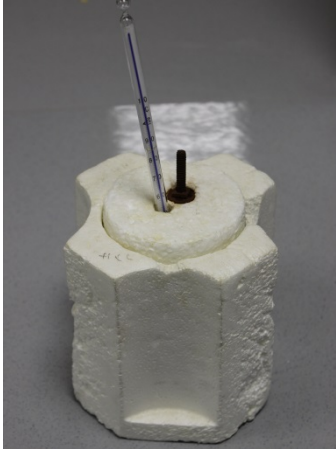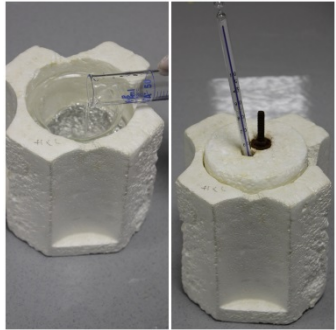
with known heats of formation of $NH_3(aq)$ and $HX(aq)$.

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 Styrofoam 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.	
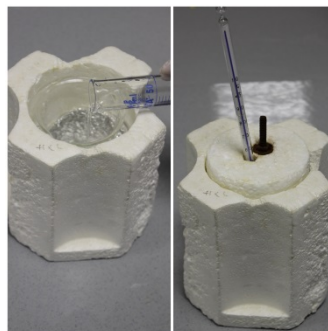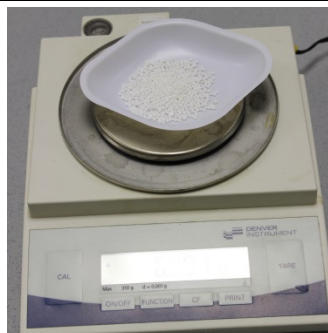
Calculations

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)]

<div align="center">**DATA SHEET**</div>

<div align="center">**Heat of Neutralization**</div>

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