

Name:

Period:

Seat#:

**Required Sections:** (Refer to R-15 for guidelines and requirements. Make note of any specific changes given by your teacher in class.)

**Prelab:** Prelab Task, Materials, Reagent Table, Procedures, and set up Data Tables before you get to class.

**During Lab:** Data section – Fill out your data table that is already set up from the prelab.

**Post-lab:** Calculation section, Discussion Questions Section (both done in lab notebook), Post-Lab Two Pager (done on separate worksheet).

## Introduction

When studying thermodynamics, the equation for free energy of a reaction,  $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ , is often encountered. In this experiment, you will use this equation to estimate the minimum entropy change required to bring about a reaction. The enthalpy change,  $\Delta H$ , and the initial temperature will be determined for a reaction during the lab. You will be able to infer/assume something conceptually which allows you to determine what  $\Delta G$  is. From these values and the equation for free energy, the minimum entropy change to bring about a *spontaneous reaction* will be estimated. \*Hint\* - think about what value of  $\Delta G$  is between a spontaneous reaction and a non-spontaneous reaction...what value is between a negative number and a positive number???

Another aspect of this lab will be trying to account for the amount of energy absorbed or lost by the calorimeter during the reaction. As the reaction proceeds, energy is being transferred between the system and the surroundings, but the surroundings include the calorimeter! The energy absorbed or lost by the calorimeter needs to be accounted for to help get a more accurate value for the reaction occurring.

## Purpose

The purpose of this experiment is to estimate the minimum entropy change required for a reaction.

**Materials** – don't forget to use an MSDS to do your reagent table!

### Chemicals

- \* Solid  $\text{NaNO}_3$
- \* Solid  $\text{NH}_4\text{Cl}$
- \* Solid  $\text{NH}_4\text{NO}_3$

### Equipment

- Thermometer probes
- Vernier interface
- Computer that has a standard USB or an adapter

- Ring stand
- Utility clamp
- Calorimeter
  - 600mL beaker
  - Styrofoam cup w/lid
- Stir bar and retriever

- Stirring plate
- 50mL graduated cylinder
- Weigh Boat
- Distilled water

## Prelab Task

Your prelab task is to write a procedure to determine the "Calorimeter Constant" for this experiment using known amounts of Hot and Cold water. Put it in the "Prelab" section of your lab notebook. Remember, flow chart format! There are lots of videos or articles online if you would like, but here is one short example video:

<https://youtu.be/ZT4U05vqHXY>



**Procedure** – Remember to make a flow chart, include diagrams/drawings of steps/equipment etc. Google "flow chart procedures" if you are not familiar with how to make a flow chart. You aren't just drawing boxes around all your sentences!

- 1) Find Calorimeter Constant first – refer to Prelab Task
- 2) Set up your calorimeter made up of a Styrofoam cup inside a 600mL beaker. A lid punched with a hole for the thermometer should also be used. Put the calorimeter on the stir plate and set the temperature probe up in the ring stand/clamp.
- 3) Place about 50 mL of distilled water in the calorimeter. Use the density of water to determine the mass of the water. Measure the temperature of the water.
- 4) Calculate the mass of solid needed to prepare 50.0 mL of a 1.00 M solution of the solid you will be using. Weigh the sample, and record the mass in the data table. Make sure to write down which solid you are using.
- 5) Add the stir bar to the calorimeter. Turn stir plate on. Make sure it's stirring gently before the next step.
- 6) Set up the temperature probe and Vernier software – **you will use the same procedures for the software that you learned in the Determining Enthalpy lab. Review that lab handout if you need a refresher on the software!** Be ready to record temperatures before you start the next step.
- 7) Add the solid to the water, and place the lid on the calorimeter, and begin collecting your temperature change with the thermometer probe. Stir gently, and record the temperature when the entire solid has dissolved.
- 8) Repeat the procedure one more times.

## Disposal and Cleanup

Your teacher will provide disposal and cleanup instructions.

## Data Table

1. Make your own data table! Remember, you need to make sure your data table has all required elements! A sample is provided below. You will need to add a descriptive title, units on all rows/columns, and a spot for qualitative data, the one below is not adequate! Remember to use enough space, make it look professional, etc!
2. Glue in a copy of your Logger Pro graph(s) below your data table.

Calorimeter Constant Determination	
Mass of cold water (g)	
Initial Temp of cold water (°C)	
Final Temp of mixed water (°C)	
Temp. change cold water (Δ°C)	
Mass of hot water (g)	
Initial Temp of hot water (g)	
Final Temp of mixed water (°C)	
Temp. change hot water (Δ°C)	

Descriptive Title	Trials	
	1	2
Solid Used:		
Mass of water (g)		
Mass of solid (g)		
Moles of solid (mol)		
Initial temperature (°C)		
Final temperature (°C)		
Temp. change (Δ°C)		

**Calculations** - Show all calculations, use proper dimensional analysis, units everywhere, proper sig figs, etc.

After performing all calculations, create a "Results Table" to record the final results of the following calculations. Same expectations as a Data Table with regards to descriptive titles, tables, units, etc. You can either average your data and then perform the calculations once, or you can do the calculations for each trial and then average your final results. Either way, be mindful of rounding issues.

**\*NOTE\*** - You do not need to gather data for the other compounds used by the other groups. We will report out to the class and discuss the results for each compound.

1. Calculate the heat capacity of the calorimeter,  $C_{cal}$  (J/°C).
  - Hint – here is a video that may help you! <https://youtu.be/QhOdtYaknA>
  - $-q_{hot} = q_{cold} + q_{cal} \rightarrow -m_{hot}C_{H_2O}\Delta T_{hot} = m_{cold}C_{H_2O}\Delta T_{cold} + C_{calorimeter}\Delta T_{cold}$
  - (using  $\Delta T_{cold}$  for the calorimeter, because it started and ended at the same temps as the cold water!)
2. Calculate the heat of the reaction.
  - Hint - The heat capacity of the calorimeter will be included in the calculation!  $q_{rxn} = -(mC\Delta T + C_{cal}\Delta T)$
3. Calculate the  $\Delta H$  for the reaction in kJ/mol using the heat of rxn and the number of moles of solid that you used.
4. Calculate the  $\Delta S$  for the reaction.



**Post Lab Discussion Questions** – Do not recopy the questions, just paraphrase them into your answer.

1. What are the units for enthalpy,  $\Delta H$ ?
2. What are the units for entropy,  $\Delta S$ ?
3. What are the units for Gibbs free energy,  $\Delta G$ ?
4. What is the algebraic sign on  $\Delta G$  for a spontaneous reaction? For a nonspontaneous reaction?
5. What can you use as the value of  $\Delta G$  for the tipping point when a reaction switches from nonspontaneous to spontaneous?
6. Write a balanced equation for the reaction your group studied (include heat written into the balanced equation as either a reactant or product).
7. Write a balanced net ionic equation for the reaction your group studied.
8. Was the reaction spontaneous? How do you know this from your observations during the lab?
9. From the temperature change of your trials, what must the sign for  $\Delta H$ ? How do you know?
10. Based on question 6/7, what must be true about the sign for  $\Delta S$ ? Explain why with support.
11. Many students believe that a reaction must be exothermic to be spontaneous. Comment on this in terms of this experiment. Looking for detailed thoughts about why someone would think a reaction must be exothermic to be spontaneous, if they are correct or not, and why.
12. Contrary to what many students would predict, the dissolution of calcium hydroxide in water has a negative entropy. Provide an explanation for why a student would wrongly assume it was a positive entropy, and why it does in fact have a negative entropy.  
$$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^{-}(aq)$$