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## Experiment #1 – Calorimetry: Heat of Solution

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When a substance undergoes a change in temperature, the quantity (Q) of heat lost or gained can be calculated using the mass (m), specific heat (s), and change in temperature ( $\Delta T = T_{\text{final}} - T_{\text{initial}}$ ):

$$Q (\text{heat}) = m s \Delta T \quad (1)$$

The specific heat (s) is the amount of heat required to change the temperature of one gram of a particular substance by one degree Celsius (or 1 K). For water, the specific heat is 1.00 calorie / g °C or 4.184 Joules / g °C.

Heat exchange experiments are conducted in a calorimeter which is an insulated container. The calorimeter is constructed such that ideally there will be no heat exchange between the contents of the calorimeter and the surrounding environment. However, heat exchanges can occur between the components within the calorimeter. This can be expressed in the following way:

$$Q_{\text{system}} = Q_{\text{gain}} + Q_{\text{loss}} \quad (2)$$

If the system does not exchange heat with the surroundings, then the  $Q_{\text{system}} = 0$ , and

$$Q_{\text{gain}} = -Q_{\text{loss}} \quad (3)$$

When a hot substance comes in contact with a cold substance within a calorimeter, heat will spontaneously pass from the hot to the cold substance until they achieve a thermal equilibrium (the same constant temperature). The quantity of heat lost and gained is equal in magnitude, but opposite in sign. For example, if a piece of hot metal is placed in cold water within a calorimeter, the following thermodynamic equation can be used:

$$Q_{\text{water}} = -Q_{\text{metal}} \quad (4)$$

Since each substance undergoes a temperature change, we can write  $Q_{\text{water}} = (m s \Delta T)_{\text{water}}$  and  $Q_{\text{metal}} = (m s \Delta T)_{\text{metal}}$ . Substituting this in the above relationship, 4, we get:

$$(m s \Delta T)_{\text{water}} = -(m s \Delta T)_{\text{metal}} \quad (5)$$

When a chemical reaction occurs in a water solution, the situation is similar to when a hot piece of metal is placed in water. When a reaction occurs, there is a change in the potential energy of the chemicals, resulting in a heat exchange in the calorimeter. As in the hot metal experiment, the heat flow of the reaction is equal in magnitude but opposite in sign to that of the solution as shown by the following equation:

$$Q_{\text{reaction}} = -Q_{\text{solution}} = -(m s \Delta T)_{\text{solution}} \quad (6)$$

By measuring the mass of the solution (the water plus the solute) and by observing the temperature change that the solution undergoes, we can calculate  $Q_{\text{solution}}$  and therefore  $Q_{\text{reaction}}$ . Because the experiment is performed under constant pressure conditions, the heat flow of the reaction is also equal to the enthalpy change,  $\Delta H$ , for the reaction. Since the mass of solute can vary for different experiments, the  $Q_{\text{reaction}}$  will also vary. Therefore,  $\Delta H$  is generally reported as the heat of reaction per a fixed amount of the reacting chemical, and we will calculate  $\Delta H$  per gram or per mole of the solute.

$$\Delta H_{\text{rxn}} = \frac{Q_{\text{rxn}}}{\text{mass}_{\text{solute}}} \quad \text{or} \quad \Delta H_{\text{rxn}} = \frac{Q_{\text{rxn}}}{\text{mole}_{\text{solute}}} \quad (7)$$

If the temperature of the water goes up, heat has been given off by the reaction, so that reaction is exothermic, and  $\Delta H$  of the reaction is negative. If the temperature of the water goes down, heat is absorbed by the reaction, therefore the reaction is endothermic, and  $\Delta H$  for the reaction is positive.

### Procedure

Construct a calorimeter, where two polystyrene coffee cups with a plastic cover on top are nested together. Weigh and record the mass of the clean, dry calorimeter. Place the calorimeter into a 400 mL glass beaker for stability during the experiment.

Place about 50 mL of distilled water in the calorimeter and weigh. Measure the temperature of the water to 0.01°C. The temperature should be within a degree or two of room temperature. Weigh out about 5 grams of the solid assigned to you. Add the solid to the calorimeter. Stirring and swirling the calorimeter, determine to 0.01°C the stable maximum or minimum temperature reached as the solid dissolves. Check to make sure that the solid is completely dissolved. A temperature change of at least five degrees should be obtained in this experiment. AFTER THE FINAL TEMPERATURE IS DETERMINED, WEIGH THE CALORIMETER WITH THE DISSOLVED SOLID.

Repeat the experiment for trial 2. If necessary, increase the amount of solid used to obtain a temperature change of at least five degrees. Use the same calorimeter containing a new 50 mL sample of water. Assume the same mass for the dry calorimeter.

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**Data and Calculations**

|   | Trial 1 | Trial 2  |
|---|---------|----------|
| Mass of empty calorimeter   | _____   | _____ g  |
| Mass of calorimeter + water                                       | _____   | _____ g  |
| Mass of calorimeter + water + solid (final mass)                  | _____   | _____ g  |
| Mass of solution  | _____   | _____ g  |
| Mass of solid (solute)  | _____   | _____ g  |
| Initial temperature   | _____   | _____ °C |
| Final temperature   | _____   | _____ °C |
| $\Delta T$  | _____   | _____ °C |
| $Q_{\text{solution}} = (m \text{ s } \Delta T)_{\text{solution}}$ | _____   | _____ J  |
| $Q_{\text{rxn}} = -Q_{\text{solution}}$                           | _____   | _____ J  |

Calculate the heat of reaction,  $\Delta H$ , per gram of solid

$$\Delta H_{\text{reaction}} = \frac{Q_{\text{rxn}}}{\text{mass}_{\text{solute}}} \quad \text{_____ J / g (trial 1)* and _____ J / g (trial 2)*}$$

\* *Note*: These values should be within 5% of each other or another trial should be done.Average  $\Delta H$  per gram \_\_\_\_\_ J / g

Molecular Formula of Solid \_\_\_\_\_ Molecular Weight \_\_\_\_\_

Calculate the molar heat of solution,  $\Delta H$ , for this solid. Use the average  $\Delta H$  value.

$$\Delta H_{\text{reaction}} \text{ _____ kJ / mol}$$

Calculate the % error for the above value.

(Literature values:  $\text{Na}_2\text{CO}_3$   $\Delta H_{\text{reaction}} = -28.1$  kJ / mol;  $\text{NH}_4\text{NO}_3$   $\Delta H_{\text{reaction}} = +25.7$  kJ / mol)

$$\text{Theoretical } \Delta H_{\text{reaction}} \text{ _____ kJ / mol} \quad \text{\% error _____}$$

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**Post-Lab Questions: Heat of Solution**

1. A metal sample weighing 63.2 g with a temperature of 100.0 °C was placed in a calorimeter containing 41.0 g of water at 24.5 °C. The equilibrium temperature of the water and metal was found to be 35.0 °C.

A. What was  $\Delta T$  for the water? ( $\Delta T = T_{\text{final}} - T_{\text{initial}}$ ) \_\_\_\_\_ °C

B. What was  $\Delta T$  for the metal? \_\_\_\_\_ °C

C. Taking the specific heat of water to be 4.184 J / g °C, calculate the specific heat of the metal using Eq. 5. \_\_\_\_\_ J / g °C

2. When 5.00 g of  $\text{KNO}_3$  were dissolved in 49.00 g  $\text{H}_2\text{O}$  at 24.00 °C inside a calorimeter, the temperature of the resulting solution fell to 15.60 °C.

A. Is this reaction endothermic or exothermic? \_\_\_\_\_  
Explain:

B. Calculate the heat lost or gained by the solution chemicals (this is  $Q_{\text{soln}}$ ) in the calorimeter.

$Q_{\text{soln}} =$  \_\_\_\_\_ Joules

C. What is  $Q_{\text{rxn}}$  for the reaction that occurred?

$Q_{\text{rxn}} =$  \_\_\_\_\_ Joules

D. Calculate the heat of this reaction,  $\Delta H$ , in Joules/g  $\text{KNO}_3$ .

$\Delta H_{\text{rxn}} =$  \_\_\_\_\_ Joules / g

E. Calculate the molar heat of this reaction,  $\Delta H$ , in kJ/mol  $\text{KNO}_3$ .

$\Delta H_{\text{rxn}} =$  \_\_\_\_\_ kJ / mole