

Chemistry 1225 Make up Lab  
ENTHALPY OF HYDRATION

All chemical reactions occurring in the world around us are occurring at essentially constant atmospheric pressure. Chemical reactions can result in release or absorption of heat resulting in a change in temperature. The thermodynamic function, enthalpy, is a measure of heat flow in processes occurring at constant pressure. Enthalpy, which is denoted by the symbol H, is a state function meaning that the enthalpy of a system depends only on the current state of the reaction not on how it got to that state. At constant pressure, the change in enthalpy is equal to the heat,  $q_p$ , gained or lost by the system (Equation 1)

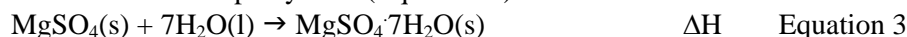
$$\Delta H = q_p \quad \text{Equation 1}$$

The heat gained or lost by a system can be calculated by measuring the temperature change in Kelvins,  $\Delta T$ , and multiplying it by the mass in grams, m, and by a quantity called specific heat, s, (Equation 2).

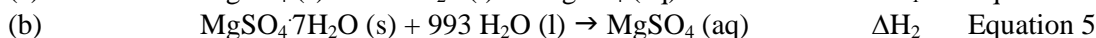
$$q_p = sm\Delta T$$

The specific heat is dependent on the chemical identity of the substance being measured. The specific heat of pure water is 4.18 J/g·K.

The enthalpy of hydration ( $\Delta H$ ) of an anhydrous salt is the enthalpy change for converting 1 mole of the anhydrous salt, A, into 1 mole of its hydrate,  $A \cdot xH_2O$ . An example of such a reaction is the conversion of magnesium sulfate into its heptahydrate (Equation 3).

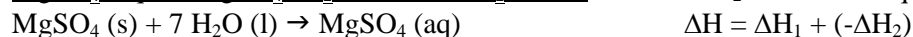
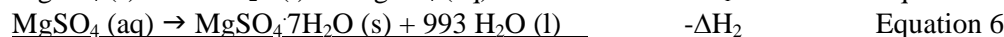


This reaction is slow so it is impossible to collect the time-temperature data to calculate the  $\Delta H$ . However it is easy to experimentally determine the enthalpy change for two separate reactions: a) dissolution of the anhydrous salt in water and b) dissolution of the hydrate in water.



Hess's law states that at constant pressure the enthalpy change for a process is not dependent on the reaction pathway but is only dependent on the initial and final states of the system. The enthalpy of individual steps in a process can be added to obtain the net enthalpy change for the overall process.

After reversing equation 5 (which results in a reversed sign of  $\Delta H$ ) and adding the resulting equation 6 to equation 4,  $\Delta H$  for the hydration process can be easily calculated. The enthalpy of hydration,  $\Delta H$ , is obtained by adding the enthalpy changes for the steps involved:  $\Delta H_1 + (-\Delta H_2)$ . Thus, to calculate  $\Delta H$  we need to calculate  $\Delta H_1$  and  $\Delta H_2$  from experimental data.



### Calculations

Show all your calculations. Use graphing paper to prepare graphs

1. a. Calculate the heat transferred during the process of dissolution of anhydrous salt. To do this, you need to know the mass of the resulting mixture, the temperature change of the resulting mixture, and the heat capacity of the mixture. Ignoring the calorimeter constant:

$$q = (\text{mass of the mixture, g}) \times (\text{specific heat of the mixture, J/g deg}) \times (\Delta T, K) \quad \text{Equation 7}$$

For this reaction,  $\Delta T$  is positive meaning that heat was released by the system and the reaction is exothermic. The heat capacity of the mixture is 3.84 J/g·K.

b. Calculate the number of moles of the salt used (n).  $n = \text{mass/molar mass}$

c.  $\Delta H_1 = -q/n$ , J/mol or kJ/mol

Equation 8

2. To calculate the  $\Delta H_2$ , the enthalpy of dissolution of hydrate, use the same equations as above (this time the  $\Delta T$  has a negative value and the reaction is endothermic).

3. Calculate  $\Delta H$ .  $\Delta H = \Delta H_1 - \Delta H_2$

### Determination of the temperature change ( $\Delta T$ )

The collected temperature data is plotted vs time and the  $\Delta T$  is determined graphically.

Although heat is transferred instantaneously at the time of mixing, thermometers exhibit a certain delayed response. Therefore it is not possible to record an accurate temperature at the point of mixing and shortly thereafter. Also, it is impossible to measure the temperature at the exact time of mixing of the reactants.

$\Delta T$  is determined by extrapolating the temperature-time line back to the point of mixing and determining the final and initial temperature of the mixture at the actual time of mixing. To prepare the graph, use graphing paper. Make sure that you draw a vertical line (line of mixing) at 5 minutes, the time of mixing. Draw the best straight line through the data points collected for pure water. Extrapolate this line to the line of mixing. The intercept with this line is the initial temperature ( $T_i$ ) of the water. Next, draw the best straight line through the data points collected after the point of mixing. The extrapolation of the line to the point of mixing will give you the best estimate of the final temperature ( $T_f$ ) of the reaction mixture (do not confuse it with the last measurement of the temperature of the mixture).

$$\Delta T = T_f - T_i$$

Equation 7

### PROCEDURE

In this experiment you will use anhydrous magnesium sulfate ( $\text{MgSO}_4$ ) and magnesium sulfate heptahydrate ( $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ , also called Epsom salt)

Caution: while handling anhydrous  $\text{MgSO}_4$  be careful to avoid having the solid absorb water from the atmosphere.

1. Tare a weighing dish.
2. Weigh between 3.50 and 4.00 g of  $\text{MgSO}_4$  in a weighing dish. Keep the salt covered until ready to use.
3. Calculate the mass of  $\text{MgSO}_4$ .
4. Prepare a Styrofoam cup (your calorimeter) and weigh it to the nearest 0.01g.
5. Add about 50 ml of distilled water to the cup.
6. Measure the temperature of the water in the cup at 30-second intervals for a period of 4.5 minutes.
7. At 5.0 minutes, quickly transfer the weighed solute ( $\text{MgSO}_4$ ) to the water in the calorimeter while stirring constantly with a glass stirring rod.
8. Measure the temperature of the mixture at 5.5 minutes, at 6.0 minutes, and at 1-minute intervals for an additional 9 minutes (a total of 15 minutes from the start of water temperature measurements).
9. Weigh the calorimeter with its contents and calculate the mass of the reaction mixture.
10. Repeat steps 1-9 using another portion of water and anhydrous salt.
11. Repeat the procedure two more times, using separate 7.65 to 7.70 g portions of  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  each time.

### Lab Report

Calculate the  $\Delta H_1$ ,  $\Delta H_2$ , and  $\Delta H$  for the hydration of magnesium sulfate. In your report tabulate your results, outline any possible sources of error, and include your graphs.

1. The following temperature vs time data were collected when 14.00 g of anhydrous sodium acetate (M.W. = 82.03 g/mol) was dissolved in 100 ml of water. The heat capacity of the reaction mixture was 4.045 J/g·K.

<u>time, min</u>	<u>temp., °C</u>	<u>time, min</u>	<u>temp., °C</u>
0.0	23.50	7.0	29.30
0.5	23.45	8.0	29.05
1.5	23.49	9.0	28.75
2.5	23.50	10.0	28.50
3.5	23.50	11.0	28.20
4.5	23.50	12.0	27.95
5.0	MIX	13.0	27.65
5.5	28.30	14.0	27.35
6.0	29.00	15.0	27.10

- Prepare a graph of the temperature vs time data.
- Determine  $\Delta T$  for the process.
- Calculate  $q_p$  (heat transferred).
- Calculate the number of moles of anhydrous sodium acetate dissolved.
- Calculate  $\Delta H_{\text{dissolution}}$  of anhydrous sodium acetate.
- Knowing that the  $\Delta H_{\text{assolution}}$  of sodium acetate trihydrate is 19.7 kJ/mol, calculate the enthalpy of hydration for anhydrous sodium acetate.

Show all your work and calculations.

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1. A student was given a sample of an anhydrous salt to determine its enthalpy of hydration. He weighed the sample but left it uncovered overnight and performed the experiment the following day.

a. Explain how the enthalpy of dissolution will be affected because of this error.

b. Would the student avoid the problems by reweighing the sample the following day, before using it for determination? Explain.

c. Explain what effect the use of a glass beaker instead of a Styrofoam cup would have on the determined  $\Delta H_{\text{hydration}}$ .

2. The student decided to check the accuracy of the thermometer and placed it in boiling water. The thermometer indicated 103.5 °C. When placed in an ice bath, the thermometer reading was 3.5 °C. Would using this thermometer in the experiment cause any errors in the calculated  $\Delta H_{\text{hydration}}$ ?