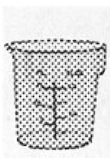
Lab 8 Enthalpy of Solution


Balance
Thermometer
Plastic weighing dish
150-mL polypropylene beaker
Wash bottle
Graduated pipette


Styrofoam cups (large 12 oz if possible) or nested paper coffee cups.
Matching size plastic cup
Distilled water

Epsom Salts

Hazards for this experiment

Epsom salts can be used as a laxative. Don’t drink large amounts of the solution.

Prelab

Reactions are performed in calorimeters to experimentally determine the ΔH_{rxn}. The class of reaction that you will be determining is the energy involved in dissolving a substance. This is called the heat of solution.

A notorious exothermic heat of solution is that of concentrated sulfuric acid. On diluting sulfuric acid, the amount of heat released is enough to boil the solution splattering hot acid over the person diluting the acid. Also, the heat is generated so quickly that the glass container can shatter from the thermal stress.

Adding water to the denser concentrated sulfuric acid will have the water layer on top of the acid creating a narrow, concentrated mixing zone that quickly erupts.

On diluting sulfuric acid, the concentrated acid should be slowly added to the water. The acid will sink through the water spreading the heating more evenly. The rule is:

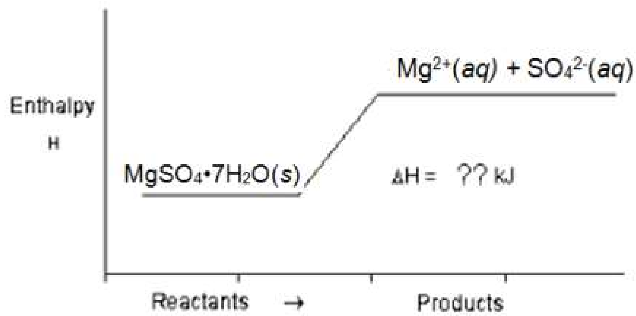
Always add acid to water as you ought’r.

The solution reaction you are going to perform is much safer than the heat of solution of sulfuric acid.

You are going to experimentally determine the $\Delta H_{solution}$ for

$MgSO_4 \bullet 7\ H_2O(s) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq) + 7\ H_2O$ (the water of hydration goes into the aqueous solution)

As you have learned, kinetic thermal energy can be absorbed as stored chemical energy. That is what this experiment is about. Magnesium sulfate solution has a greater enthalpy of formation than hydrated magnesium sulfate solid. The change in stored energy, ΔH_{rx}, is positive resulting in an endothermic reaction.



The graph is a common way of expressing the enthalpy changes in a reaction. Note that the stored chemical energy, *H*, is the *y* axis value and the progress of the reaction from reactants to products is reflected in the *x* axis. Any endothermic reaction is an increase in enthalpy. The difference is the ΔH_{rx} of the reaction. While the value of the difference can be measured experimentally, the absolute amount of enthalpy cannot be measured.

Procedure:

1. Assemble your calorimeter by nesting two Styrofoam cups together and then put in a plastic cup liner. The plastic cup liner will provide some protection from puncturing the Styrofoam cup when you stir the solution.

2. Find the mass of the dry calorimeter without the lid and enter the data in your lab booklet.

3. Pour approximately 50 mL of water into your calorimeter. Let it stand for at least an hour to be sure that it has acclimatized to the room temperature.

4. Fill your wash bottle with water and let it stand next to your calorimeter so that the water in it will have the same temperature as the water in the calorimeter.

5. Get your package of Epsom salts

6. Zero (tare) the clean, dry 150-ml propylene beaker on the balance. Take the beaker off the balance and add about two tablespoons of the Epsom salts to the beaker. Put the beaker and Epsom salt on the balance and record the mass. If you don't have at least 16 g of Epsom salts, add another half tablespoon and reweigh. Proper weighing technique recommends that the solids are added to the weigh container while not on the balance so that chemicals don't have a chance of getting on the balance.

7. Record the temperature of the water in the calorimeter, T_1. Estimate to the 0.1°C.

8. Pour the Epsom salts into the calorimeter containing the water.

9. Using your wash bottle, rinse any remaining solute left in the beaker into the calorimeter.

10. Stir vigorously with the thermometer, taking care not to puncture the calorimeter, until the solution stops dropping in temperature (about 3 minutes). Once the temp has stopped dropping, you can assume that the Epsom salts has completely dissolved.

11. Take the lowest temperature of the solution estimating to the nearest tenth of a degree. Don't empty the calorimeter yet. You must find the mass of the solution

12. Carefully find the mass of the calorimeter again. Ionic solutions are fatal to electronics so take care not to spill!

AP L08	*Enthalpy of Solution*	*Unambiguous Date*

02

Purpose: *To find the enthalpy of solution of Epsom salts*

Equipment:

Balance	*Thermometer*	*Styrofoam cup calorimeter*
50-mL beaker	*Wash bottle*	

Procedure:

You should be able to write your own procedure now, especially if you think you will need this notebook for college credit later. This is a good practice. Think of it as a diary of your experiment. Write it as you do your experiment. Remember that a lab notebook is not expected to be perfect because it is done as you do experiments. Just cross out any error with a single line. No erasing or white out.

Data:

(a) Mass of dry, empty calorimeter: *4.10 g*

(b) Mass of the Epsom salts *16.36 g*

(c) Water in calorimeter: T_1 *= 19.2 ºC* *Solution:* T_2 *= 13.4 ºC*

(d) Mass of solution and calorimeter: *59.21 g*

Calculations:

(d) – (a) Mass of solution: 55.11 g

$\Delta T_{soltuion} = T_2 - T_1 = -5.8\ ºC$

s= specific heat of solution[1]*= 4.18 J/g°C.*

Energy absorbed by solution in calorimeter, $q_{calorimeter} = m_{sol'nr} \times c_{sol;n} \times \Delta T_{solution}$

$q_{calorimeter} = 55.11\ g \times 4.18\ J/g°C \times (-5.8\ ºC)$

$q_{calorimeter} = -1300\ J$

[1] For dilute solutions the specific heat can be assumed to be that of water.

The solution's kinetic energy was changed into potential chemical energy as the Epsom salts dissolved.

The Epsom salt change in energy (gain) was opposite the solution's (lost)

$$q_{reaction}, = - q_{calorimeter}$$

$$q_{reaction} = +1300 \text{ J}$$

To convert the $q_{reaction}$ to ΔH_{rx} the moles of the reactant must be found.

Assume that the reactant was pure Epsom salts which is magnesium sulfate heptahydrate, $MgSO_4 \cdot 7\ H_2O$

molar mass = 246.5 g/mol

Calculation to find mol of $MgSO_4 \cdot 7\ H_2O$

$$16.36 \text{ g } MgSO_4 \text{ g } 7\ H_2O \times \frac{1 \text{ mol}}{246.5 \text{ g}} = 0.06637 \text{ mol } MgSO_4 \text{ g } 7\ H_2O$$

$$\Delta H_{rx} = q_{rx} / \text{mol } MgSO_4 \cdot 7\ H_2O$$

$$\Delta H_{rx} = 20. \text{ kJ / mol } \ MgSO_4 \cdot 7\ H_2O\ (s) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq)$$

The published values for this solution reaction assume infinite dilution, $\Delta H_{solution}$ = 16 kJ/mol

Calculate your % error based on this value. *25% error*

Considering that the temp change had only 2 sig figs, the best that could be expected would be a 10% error. A larger value than expected indicates that either there was more Epsom salts than I measured or that the temperature change was greater than expected. The weighing error doesn't seem likely, but a lower temperature reading could have resulted from incomplete mixing. The salt was hard to dissolve and my thermometer could have been in an area where the Epsom salt was more concentrated resulting in a greater lowering of temperature. As the temperature change was small, even a 1°C error could have led to a 20% error.