ENTROPY OF DISSOLVING

Introduction:
When a substance dissolves, heat is usually absorbed or released. For example, in the standardization of NaOH lab, the dissolving of the NaOH in water was very exothermic, so you used an ice bath. This heat is called the “heat of solution.” Energy is gained or lost by the solvent and the surroundings in the process. Therefore we can use calorimetry to determine the heat of solution of a solid/solvent pair. It is easy to confuse the heat of solution with the heat gained or lost by a solution, \( q_{\text{soln}} \). You have to consider the context to know which one is being discussed.

In this experiment, you will use the Gibbs free energy equation, \( \Delta G = \Delta H - T \Delta S \), to estimate the minimum entropy change required to create a solution.

Purpose:
The purpose of this experiment is to estimate the minimum entropy change required for a spontaneous process.

AP Chemistry Curriculum Alignment

Learning Objective 5.7: The student is able to design and/or interpret the results of an experiment in which calorimetry is used to determine the change in enthalpy of a chemical process (heating/cooling, phase transition, or chemical reaction) at constant pressure. [See SP 4.2, 5.1, 6.4]

Learning Objective 5.6: The student is able to use calculations or estimations to relate energy changes associated with heating/cooling a substance to the heat capacity, relate energy changes associated with a phase transition to the enthalpy of fusion/vaporization, relate energy changes associated with a chemical reaction to the enthalpy of the reaction, and relate energy changes to P\( \Delta V \) work. [See SP 2.2, 2.3]

Equipment/Materials:
solid samples: NaNO\(_3\), NH\(_4\)Cl, NH\(_4\)NO\(_3\)  thermometer  weigh boats, spatulas

           calorimeter  rubber policeman (opt.)

           distilled water  analytical or triple-beam balance

Safety:
- An apron and goggles must be worn at all times in the lab.
- Sodium nitrate is hazardous in case of ingestion. Slightly hazardous in case of skin contact (irritant), eye contact (irritant), or inhalation. Prolonged exposure may result in skin burns and ulcerations. Over-exposure by inhalation may cause respiratory irritation.
- Ammonium chloride is hazardous in case of eye contact (irritant). Slightly hazardous in case of skin contact (irritant, sensitizer), ingestion, or inhalation.
- Ammonium nitrate is a strong oxidizer. Contact with other material may cause a fire. May cause eye and skin irritation. Causes respiratory tract irritation.

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Prelab:

1. Choose one of the solids listed above.
   
   a. Calculate the mass of your solid needed to prepare 10.0 mL of a 1.00 M solution of the solid you will be using. Show your work.
   
   Solid chosen ____________________  Mass needed ____________

   b. Write the dissociation equation for your chosen solid in water. Show states.

2. A student dissolved some urea in water and calculated the ΔH of solution to be 14 kJ/mol at 25°C. What is the minimum S for this process? (Hint: it was spontaneous.)

Procedure:

1. Obtain a small black plastic calorimeter and thermometer (or LabQuest and temperature probe). Weigh and record the mass of the calorimeter.

2. Fill the calorimeter about ¾ full with distilled water. Reweigh and record.

3. Measure the (initial) temperature of the water. Record.

4. Weigh the solid you will be using, and record its actual mass. It does not need to be exactly what you calculated in the PreLab.

5. Hold the calorimeter near the top so that you don’t heat it with your hand. Add the solid to the water. (Hint: the weigh boats fold!) Stir gently, and record the lowest temperature reached. All of the solid should be dissolved.

6. Dispose of the solution in the sink. Rinse and dry the calorimeter.

7. Repeat the procedure two more times.

Data:

<table>
<thead>
<tr>
<th>Solid used: ____________________</th>
<th>Object</th>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Trial 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass, g</td>
<td>calorimeter</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>water and calorimeter</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td></td>
<td>weigh boat</td>
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<tr>
<td></td>
<td>weigh boat and solid</td>
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<td></td>
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<tr>
<td>Temperature, °C</td>
<td>Initial</td>
<td></td>
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<tr>
<td></td>
<td>Final</td>
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</tbody>
</table>
Calculations/Analysis: Show all work with proper sig figs and units.

1. Calculate the heat of solution \( q_{\text{rxn}} \) for each trial. Find \( q_{\text{soln}} \) first. The heat capacity of the calorimeter will not be included in the calculation. Assume the specific heat capacity of the solution to be that of water.

2. Calculate the \( \Delta H \) for each trial in kJ/mol using the heat of solution \( q_{\text{rxn}} \) and the moles of the solid.

3. Average the \( \Delta H \)s for your trials.

4. Use the Gibbs free energy equation to calculate the minimum \( \Delta S \) for the dissolving of your solid. (Remember that \( \Delta G \) must be <0 for a spontaneous process.)

Results:

<table>
<thead>
<tr>
<th></th>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Trial 3</th>
<th>Average</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass water, g</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Moles of solid, mol</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Temp. change, °C</td>
<td></td>
<td></td>
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<tr>
<td>Heat of reaction, J</td>
<td></td>
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<td></td>
</tr>
<tr>
<td>( \Delta H ), kJ/mole</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Minimum ( \Delta S ), J/mol*K</td>
<td></td>
<td></td>
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</tbody>
</table>

Post-Lab Questions:

1. Was the dissolution spontaneous? How do you know this?
2. What does your answer to question 1 mean about the sign of \( \Delta G \)? Explain.
3. From the temperature change of your trials, what must be the sign for \( \Delta H \)? Explain.
4. Considering your answers to questions 2 & 3, what must be true about the sign for \( \Delta S \)? Explain.
5. Write a balanced thermochemical equation for the dissolution of the substance you studied, based on your results.
6. Many students believe that a reaction must be exothermic to be spontaneous. Using your results and CER, support or refute this idea.

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7. Describe two sources of error in this experiment and how each of them would have affected your data and results numerically. Be as specific as possible on the effects on each variable involved in the calculations.

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