**CHM111 Lab – Enthalpy of Hydration of Sodium Acetate – Grading Rubric**

<table>
<thead>
<tr>
<th>Criteria</th>
<th>Points possible</th>
<th>Points earned</th>
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<tbody>
<tr>
<td><strong>Lab Performance</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Printed lab handout and rubric was brought to lab</td>
<td>3</td>
<td></td>
</tr>
<tr>
<td>Safety and proper waste disposal procedures observed</td>
<td>2</td>
<td></td>
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<tr>
<td>Followed procedure correctly without depending too much on instructor or lab partner</td>
<td>3</td>
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<tr>
<td>Work space and glassware was cleaned up</td>
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<tr>
<td><strong>Lab Report</strong></td>
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<tr>
<td>Data recorded clearly with proper units</td>
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<td></td>
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<tr>
<td>Calculations for q shown clearly with correct signs</td>
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<td></td>
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<tr>
<td>Question 1</td>
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<td>Question 2</td>
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<tr>
<td>Question 3 – Hess’s Law calculation shown and explained clearly</td>
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<td></td>
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<tr>
<td>Question 4</td>
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<td>Question 5</td>
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<tr>
<td><strong>Total</strong></td>
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Subject to additional penalties as per the instructor
Enthalpy of Hydration of Sodium Acetate

**Goal**
To determine the enthalpy (ΔH) for the following process:

\[
\text{NaC}_2\text{H}_3\text{O}_2(\text{s}) + 3 \text{H}_2\text{O}(\text{l}) \rightarrow \text{NaC}_2\text{H}_3\text{O}_2\cdot3\text{H}_2\text{O}(\text{s})
\]

**Introduction**
Most chemical reactions are accompanied by a change in energy (heat). The reactions will either give off heat or absorb heat. Thermochemistry is the study of heat exchanged during reactions. In thermochemistry, the atoms and molecules doing the reacting are referred to as the system and everything else in the universe is referred to as the surroundings. Reactions that release heat to the surroundings are exothermic reactions – these reactions cause the surroundings to increase in temperature. A reaction that absorbs heat is an endothermic reaction and it causes the surroundings to decrease in temperature.

The sign for the change in enthalpy of a reaction (ΔH) is negative for exothermic reactions and positive for endothermic reactions. Consider the following exothermic reaction:

(1) \[ \text{NaOH}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \quad \Delta H = -44.4 \text{ kJ/mol} \]

When one mole of sodium hydroxide is dissolved in water, the reaction (the system) releases 44.4 kJ of energy. This energy is transferred to the surroundings and the temperature of the surroundings will increase. ΔH experiments are often done in aqueous solution in insulated coffee cups. The coffee cup ensures that virtually all of energy in the system is transferred to the solution, not to the lab bench or the air.

**Calculating q_{reaction}**
The first law of thermodynamics states that the net change in energy in the universe is equal to zero:

(2) \[ \Delta E_{\text{universe}} = q_{\text{system}} + q_{\text{surroundings}} = 0 \]

Rearranging equation (2) and noting that in this case \( q_{\text{system}} \) is equivalent to \( q_{\text{reaction}} \), and \( q_{\text{surroundings}} \) is equivalent to \( q_{\text{solution}} \), we get:

(3) \[ q_{\text{system}} = -q_{\text{surroundings}} \quad \text{or} \quad q_{\text{reaction}} = -q_{\text{solution}} \]

Heat for the system and surroundings have the same magnitude, but opposite sign. Just like ΔH, when heat is lost, q is negative, when heat is gained, q is positive.

\( q_{\text{reaction}} \) cannot be determined directly. However, \( q_{\text{solution}} \) can be determined using the specific heat equation (4). m is the total mass of the solution, \( C_s \) is the specific heat of the solution (usually similar to the specific heat of water for dilute solutions) and \( T_f \) and \( T_i \) are the final and initial temperatures of the solution.

(4) \[ q_{\text{solution}} = m \ C_s \ (T_f - T_i) \]

Note that according to equation (4), when the temperature of the solution increases, the sign of \( q_{\text{solution}} \) will be positive, when the temperature of the solution decreases, the sign of \( q_{\text{solution}} \) will be negative.

**Calculating ΔH**
ΔH is the heat released when the amounts written in the balanced equation are reacted. For equation (1), that would be 1 mole of NaOH (s) reacting. However, reacting an entire mole is wasteful and unnecessary. To find ΔH, we react a much smaller amount, find the heat released or absorbed (\( q_{\text{reaction}} \)), then extrapolate how much heat would have been released if we reacted one mole. ΔH is calculated as follows:

(5) \[ \Delta H = \frac{q_{\text{reaction}}}{\text{mol}} \]
For an example of how the $\Delta H$ in equation (1) might have been determined, if 0.013 moles of NaOH are dissolved, $-572 \text{ J}$ of heat will be released. Plugging into equation (5) $\Delta H = -572 \text{ J}/0.013 \text{ mol} = -44,000 \text{ J} = 44 \text{ kJ}$. In this way we are able to find $\Delta H$ using much smaller quantities of reagents.

In this experiment you will be calculating the $\Delta H$ for dissolving anhydrous sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2 \text{(s)}$ (equation 6) and $\Delta H$ for dissolving sodium acetate trihydrate, $\text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O} \text{(s)}$ (equation 7).

$$\text{(6)} \quad \text{NaC}_2\text{H}_3\text{O}_2 \text{(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \quad \Delta H_{\text{reaction}} = ? \quad \text{(for the anhydrous)}$$

$$\text{(7)} \quad \text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O} \text{(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + 3 \text{H}_2\text{O(}l\text{)} \quad \Delta H_{\text{reaction}} = ? \quad \text{(for the hydrate)}$$

$\Delta H$ relationships and Hess’ Law

Ultimately in this lab, we want to calculate the $\Delta H$ for anhydrous sodium acetate converting to sodium acetate trihydrate (equation 8). However, measuring this directly in the lab is not feasible – the process cannot be done in solution, it will occur slowly over several days if anhydrous sodium acetate is exposed to atmospheric moisture (left in an open container).

$$\text{(8)} \quad \text{NaC}_2\text{H}_3\text{O}_2 \text{(s)} + 3 \text{H}_2\text{O(}l\text{)} \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O} \text{(s)} \quad \Delta H_{\text{reaction}} = ?$$

Since we cannot figure out $\Delta H$ for equation (8) directly, we will use Hess’s Law to determine it indirectly. Hess’ Law works because enthalpy is a state function – it is independent of how the system is prepared or the path of the reaction. Hess’ Law states that the change in enthalpy of a stepwise reaction is equal to the sum of the changes in enthalpy for the individual steps. We may not be able to measure the enthalpy for $A + B \rightarrow D$ or equation (8) but we can calculate it if we know the enthalpies of the other steps.

$$A + B \rightarrow C \quad \Delta H = -100 \text{ kJ}$$
$$C \rightarrow D \quad \Delta H = +25 \text{ kJ}$$
$$A + B \rightarrow D \quad \Delta H = -75 \text{ kJ}$$

Note that equations (6) and (7) do not add to equation (8) as written. You will need to manipulate the reactions so that they add to the reaction of interest. Other $\Delta H$ relationships are that reversing a reaction reverses the sign of $\Delta H$, and when multiplying reaction coefficients by a factor, $\Delta H$ is multiplied by the same factor.

If $A + B \rightarrow C \quad \Delta H = -100 \text{ kJ}$, then $C \rightarrow A + B \quad \Delta H = +100 \text{ kJ}$

If $C \rightarrow D \quad \Delta H = +25 \text{ kJ}$, then $2 C \rightarrow 2 D \quad \Delta H = +50 \text{ kJ}$

You will use these principles to make equations (6) and (7) add to the overall equation (8). The sum of the transformed $\Delta H$’s will then equal the $\Delta H$ of the overall reaction (8).
Laboratory Activity

**Equipment**
- 1 Large test tube
- Thermometer

**Chemicals**
- Anhydrous sodium acetate
- Sodium acetate trihydrate

A. Dissolution of Anhydrous Sodium Acetate

1. Obtain a large dry test tube and a thermometer.

2. Measure 10.0 mL of deionized water in a graduated cylinder. Record the temperature of the deionized water.

3. Weigh between 1.20 to 1.25 grams of anhydrous sodium acetate (NaC₂H₃O₂(s)). Record the exact mass. **Close the lid of the jar immediately after weighing** – anhydrous sodium acetate will absorb atmospheric moisture and convert to the hydrate if it is exposed to air – we want to keep it anhydrous!

4. Working quickly, place the solid into the test tube. Hold the test tube at the top or in a rack to avoid transferring heat from your hands. Add the 10.0 mL of water to the test tube. Be sure that there are no solids sticking to the sides. Stir gently with your thermometer. Measure the solution’s temperature as you stir, and record the highest temperature that the solution reaches.

5. Pour the solution down the sink and wash dry the test tube.

B. Dissolution of Sodium Acetate Trihydrate

6. Measure another 10.0 mL of water. Record the temperature of the deionized water.

7. Weigh out between 1.96 to 2.04 grams of sodium acetate trihydrate (NaC₂H₃O₂·3H₂O(s)). Record the exact mass.

8. Place the solid in the washed test tube and dried test tube. Add the 10.0 mL of water, and begin to stir gently. Again, be sure that no solid is stuck to the sides of the test tube. Measure the solution’s temperature as you stir, and record the lowest temperature that the solution gets to.

9. Pour the solution down the sink and wash the test tube.

C. Calculations

10. For both Part I and Part II, calculate the heat gained or lost by the solution using equation (3). Assume the specific heat of the solution is the same as pure water (4.18 J/g °C) and that the solution has the same density as pure water (1.00 g/mL).

   The heat gained or lost by the reaction should be of the same magnitude as heat gained or lost by the reaction, but have the opposite sign. Use equation (2) to find q reaction.

   Find ΔH for equations (5) and (6) by dividing q_react by moles of reactant (equation 4)

   Finally, manipulate equations (5) and (6) so that they add to equation (7) and apply Hess’s Law to determine its ΔH.
Show all your work clearly and include appropriate units.

**A: Dissolution of Anhydrous Sodium Acetate**

<table>
<thead>
<tr>
<th>Mass of water</th>
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</thead>
<tbody>
<tr>
<td>Mass of NaC₂H₃O₂</td>
<td></td>
</tr>
<tr>
<td>Mass of solution</td>
<td></td>
</tr>
<tr>
<td>Room Temperature</td>
<td></td>
</tr>
<tr>
<td>Final Temperature</td>
<td></td>
</tr>
<tr>
<td>ΔT</td>
<td></td>
</tr>
<tr>
<td>Specific heat capacity of the solution</td>
<td>4.18 J/g°C</td>
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</tbody>
</table>

Calculations for $q_{solution}$:

$q_{solution} = \text{________________}_{}$  
$q_{reaction} = \text{________________}_{}$

**B. Dissolution of Sodium Acetate Trihydrate**

<table>
<thead>
<tr>
<th>Mass of water</th>
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</tr>
</thead>
<tbody>
<tr>
<td>Mass of NaC₂H₃O₂·3H₂O</td>
<td></td>
</tr>
<tr>
<td>Mass of solution</td>
<td></td>
</tr>
<tr>
<td>Room Temperature</td>
<td></td>
</tr>
<tr>
<td>Final Temperature</td>
<td></td>
</tr>
<tr>
<td>ΔT</td>
<td></td>
</tr>
<tr>
<td>Specific heat capacity of the solution</td>
<td>4.18 J/g°C</td>
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</tbody>
</table>

Calculations for $q_{solution}$:

$q_{solution} = \text{________________}_{}$  
$q_{reaction} = \text{________________}_{}$
Show all your work and include appropriate units.

Q1. Calculate moles of NaC$_2$H$_3$O$_2$ (s), then use $q_{\text{reaction}}$ to calculate $\Delta H_{\text{reaction}}$ for anhydrous sodium acetate in kilojoules.

\[ \text{NaC}_2\text{H}_3\text{O}_2 (s) \rightarrow \text{Na}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) \]

Q2. Calculate moles of NaC$_2$H$_3$O$_2$·3H$_2$O (s), then use $q_{\text{reaction}}$ to calculate $\Delta H_{\text{reaction}}$ for sodium acetate trihydrate in kilojoules.

\[ \text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O} (s) \rightarrow \text{Na}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) + 3 \text{H}_2\text{O}(l) \]

Q3. Use Hess’ Law and the results from questions 1 and 2 to find the enthalpy of hydration ($\Delta H_{\text{reaction}}$) for anhydrous sodium acetate in kJ. Show your work clearly – rewrite transformed reaction steps from questions 1 and 2 above.

\[ \text{NaC}_2\text{H}_3\text{O}_2(s) + 3 \text{H}_2\text{O}(l) \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O}(s) \]
Q4. Another way to calculate $\Delta H$ values that cannot be measured experimentally is with standard heats of formation.

a) Using the $\Delta H^\circ_f$ values provided, calculate $\Delta H_{rxn}^\circ$ for $\text{NaC}_2\text{H}_3\text{O}_2(s) + 3 \text{H}_2\text{O}(l) \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O}(s)$ with the equation:

$$\Delta H_{rxn}^\circ = \Sigma n_p \Delta H^\circ_f(\text{products}) - \Sigma n_r \Delta H^\circ_f(\text{reactants})$$

<table>
<thead>
<tr>
<th></th>
<th>$\Delta H^\circ_f$ (kJ/mol)</th>
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<tbody>
<tr>
<td>$\text{NaC}_2\text{H}_3\text{O}_2(s)$</td>
<td>-708.8</td>
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<tr>
<td>$\text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O}(s)$</td>
<td>-1604</td>
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<tr>
<td>$\text{H}_2\text{O}(l)$</td>
<td>-285.8</td>
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b) Compare your value from 4a) to your experimentally determined value in Q3. Determine the percent error.

Q5. Use the standard reaction enthalpies given below to determine $\Delta H^\circ_{rxn}$ for the following reaction. Show your work in detail.

Find: $2 \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CH}_2\text{CO} + 3 \text{H}_2\text{O}$  $\Delta H^\circ_{rxn} =$ ?

Given:
- $\text{CH}_2\text{CO} + 2 \text{O}_2 \rightarrow 2 \text{CO}_2 + \text{H}_2\text{O}$  $\Delta H^\circ_{rxn} = -981.1 \text{ kJ}$
- $\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$  $\Delta H^\circ_{rxn} = -802.3 \text{ kJ}$