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## Carolina Investigations™ for AP Chemistry Fundamentals of Calorimetry

### Guided Activity

#### Background

Have you ever been injured and been given an instant cold pack or hot pack to place on the injury? These packs usually function by mixing water and a salt. Squeezing the pack breaks an inner packet, releasing water, which then mixes with the salt and initiates the temperature change. Depending on the salt contained in the pack, the temperature may increase or decrease. How can you determine how much heat is absorbed or released when one of these packs is used?

Calorimetry is the science of measuring heat. Many chemical and physical transformations involve energy transfer in the form of heat. The magnitude of the heat and the direction of transfer may be determined by use of a calorimeter. In reactions that occur in aqueous solutions, the energy is transferred to or taken away from the water. A calorimeter is an apparatus that is insulated and prevents heat from flowing into or out of the system. Since the experiment is run under constant pressure (atmospheric), the change in temperature of the water that is measured is due to the enthalpy of reaction. The heat of reaction may be calculated using the values measured for change in temperature of the water.

The heat transfer or change in enthalpy in a reaction ( $q_{\text{reaction}}$ ) is related to the mass of the solution ( $m$ ), the specific heat capacity of the solution ( $c$ ), and the temperature change ( $\Delta T = T_{\text{final}} - T_{\text{initial}}$ ) as described in the following equation:

$$q_{\text{reaction}} = -(m \times c \times \Delta T)$$

The specific heat capacity of a substance is the amount of energy required to raise 1 gram of the substance 1 degree Celsius. The specific heat capacity of water is  $4.186 \text{ J}/(\text{°C} \times \text{g})$ . In experiments conducted in solution, the specific heat capacity of water is generally used.

In calculating heat, an adjustment must be made for any energy transfers to or from the calorimeter. The heat capacity of the calorimeter,  $C$ , is measured in joules per degree Celsius ( $\text{J}/\text{°C}$ ). To adjust the heat of reaction for the heat absorbed or released by the calorimeter must be added to the heat absorbed or released by the solution. The change in temperature of the calorimeter is assumed to be the same as the change in temperature of the solution.

$$q_{\text{reaction}} = -[(m_{\text{water}} \times c_{\text{water}} \times \Delta T_{\text{water}}) + (C_{\text{calorimeter}} \times \Delta T_{\text{calorimeter}})]$$

The enthalpy change in the reaction,  $\Delta H_{\text{rxn}}$ , is a measure of the heat flow and at constant pressure is equal to heat flow,  $q_{\text{reaction}}$ . Enthalpy change, however, is generally written as a molar quantity. To calculate the molar enthalpy of a reaction, divide the amount of heat in the reaction by the number of moles of reactant or product ( $\Delta H_{\text{reaction}} = q_{\text{reaction}}/\text{mol}$ ). If a process increases the solution temperature it is exothermic (negative  $\Delta H$ ), whereas if a process decreases the solution temperature it is endothermic (positive  $\Delta H$ ).



4. Measure the temperature of the water in the calorimeter and record it (to  $\pm 0.1^\circ\text{C}$ ) as the initial temperature of cold water.
5. Place 100.0 mL water in a clean, dry beaker.
6. Heat the water in the beaker until it is at least  $30^\circ\text{C}$  warmer than the cold water.
7. Pour 50.0 mL of this warm water into the remaining foam cup (not part of the calorimeter).
8. Measure the temperature of the warm water and record it (to  $\pm 0.1^\circ\text{C}$ ) as the initial temperature of the warm water.
9. One group member must begin to monitor time immediately as the water is combined in the next step.
10. Remove the lid and pour the warm water into the cold water in the calorimeter. Place the lid and thermometer on the calorimeter.
11. On the Data Table, record the first temperature 10 seconds after combining the two water samples and then every 10 seconds for 90 seconds.
12. Discard the water and thoroughly dry the calorimeter.

*Determining the enthalpy of solution for calcium chloride*

1. Measure 100.0 mL water with a graduated cylinder and pour it into the calorimeter.
2. Place the lid on the calorimeter and open the tab to insert a thermometer.
3. Stir the water with the thermometer, monitoring the temperature until it is stable. Record this temperature (to  $\pm 0.1^\circ\text{C}$ ) in the Data Table as the initial temperature.
4. Measure 5.00 g calcium chloride ( $\text{CaCl}_2$ ). Record the exact value used in the Data Table.
5. Remove the lid and add the 5.00 g  $\text{CaCl}_2$  to the calorimeter and stir. Replace the lid and thermometer.
6. Continue to stir and monitor the temperature for 2 minutes. Record the highest temperature obtained (to  $\pm 0.1^\circ\text{C}$ ) in the Data Table as the final temperature.
7. Discard the solution as directed by your teacher and rinse the inner cup. Thoroughly dry the calorimeter apparatus before reusing.
8. Repeat steps 1–8 with 10.00 g  $\text{CaCl}_2$  and 15.00 g  $\text{CaCl}_2$ .

*Determining the enthalpy of solution for ammonium chloride*

1. Measure 100.0 mL water with a graduated cylinder and pour it into the calorimeter.
2. Place the lid on the calorimeter, pulling back the tab to form an opening for a thermometer.
3. Stir the water with the thermometer, monitoring the temperature until it is stable.
4. Measure 5.00 g ammonium chloride ( $\text{NH}_4\text{Cl}$ ) (recording the precise value in the Data Table) and add it to the calorimeter and stir.
5. Continue to stir and monitor the temperature for 2 minutes. Record the lowest temperature obtained (to  $\pm 0.1^\circ\text{C}$ ) in the Data Table as the final temperature.
6. Discard the solution as directed by your teacher, and thoroughly rinse and dry the calorimeter.
7. Repeat steps 1–6 with 10.0 g  $\text{NH}_4\text{Cl}$  and then with 15.0 g.

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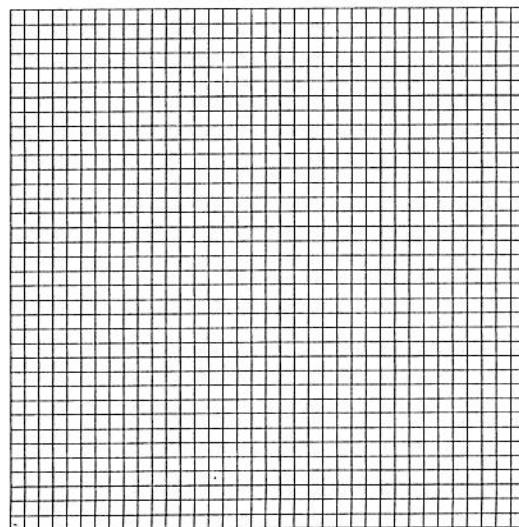
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### Data Tables

*Determining the specific heat capacity of the calorimeter*

Time (seconds)	Temperature (°C)
10	
20	
30	
40	
50	
60	
70	
80	
90	

Initial temperature of cold water (°C)	
Initial temperature of warm water (°C)	
Average temperature (°C)	
Temperature at 0 seconds (°C)	
Heat lost by the water (J)	
Heat capacity of the calorimeter (J/°C)	



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### Data Tables

*Determining the heat of solution of calcium chloride and ammonium chloride*

	Calcium chloride			Ammonium chloride		
	5 g CaCl <sub>2</sub>	10 g CaCl <sub>2</sub>	15 g CaCl <sub>2</sub>	5 g NH <sub>4</sub> Cl	10 g NH <sub>4</sub> Cl	15 g NH <sub>4</sub> Cl
Mass of water (g)						
Mass of salt (g)						
Moles of salt (g × mol/g)						
Initial temperature (°C)						
Final temperature (°C)						
Change in temperature (°C)						
Heat absorbed by the solution (J)						
Heat capacity of the calorimeter (J/°C)						
Heat absorbed by the calorimeter (J)						
Enthalpy of solution (J)						
Enthalpy of solution (kJ)						
Enthalpy/mole of solution (kJ/mol)						

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## Data Analysis

### *Determining the specific heat capacity of the calorimeter*

1. Plot a graph of the temperature compared to time. Leave a space for temperature at time zero.
2. Draw a line of best fit for the data points.
3. Based on the line, determine the temperature at a time of zero seconds ( $T_0$ ). This is the instantaneous temperature at mixing.
4. Calculate the average temperature ( $T_a$ ), using the initial temperature of the hot and cold water.
5. The difference between the average temperature ( $T_a$ ) and the calculated temperature at zero seconds ( $T_0$ ) is due to the heat lost by the water, which is based on the heat capacity of the calorimeter. The heat lost by the water is equal to the heat gained by the calorimeter. Calculate the heat lost by the water.

$$q_w = c_w \times m_w \times (T_0 - T_a)$$

6. Calculate the heat capacity ( $C$ ) of the calorimeter based on the amount of heat ( $q$ ) that the calorimeter absorbs for each change in temperature by  $1^\circ\text{C}$ . The change in temperature is the difference between the temperature of the water at zero seconds ( $T_0$ ) and the initial temperature of the cold water ( $T_c$ ).

$$C = \frac{q_w}{(T_0 - T_c)}$$

### *Determining the enthalpy of solution for calcium chloride*

1. Determine the change in temperature.

$$\Delta T = T_f - T_i$$

2. Calculate the heat absorbed by the water for each mass of  $\text{CaCl}_2$  and the heat absorbed by the calorimeter for each mass of  $\text{CaCl}_2$ . The specific heat of water is  $4.186 \text{ J}^\circ\text{C} \times \text{g}$ .

$$q_w = -[c_w \times m_w \times \Delta T] \quad q_c = -C \times \Delta T$$

3. Calculate the enthalpy of the solution for each mass of  $\text{CaCl}_2$ . The enthalpy of the solution is the heat absorbed by the water plus the heat absorbed by the calorimeter.
4. Indicate whether the process was exothermic or endothermic.
5. For each mass, calculate the enthalpy per mole of calcium chloride.
6. Graph the change in temperature of calcium chloride compared to the mass of calcium chloride. Draw a line of best fit for the points.

### *Determining the enthalpy of solution for ammonium chloride*

1. Determine the change in temperature. Record this value on the Data Sheet.

$$\Delta T = T_f - T_i$$

2. Calculate the heat absorbed by the water for each mass of  $\text{NH}_4\text{Cl}$  and the heat absorbed by the calorimeter for each mass of  $\text{NH}_4\text{Cl}$ . The specific heat of water is  $4.186 \text{ J}^\circ\text{C} \times \text{g}$ . Record these values on the Data Sheet.

$$q_w = -[c_w \times m_w \times \Delta T] \quad q_c = -C \times \Delta T$$

3. Calculate the enthalpy of the solution for each mass of  $\text{NH}_4\text{Cl}$ . The enthalpy of the solution is the heat absorbed by the water plus the heat absorbed by the calorimeter. Record these values on the Data Sheet.
4. Indicate whether the process was exothermic or endothermic.
5. For each mass, calculate the enthalpy per mole of ammonium chloride. Record these values on the Data Sheet.
6. Graph the change in temperature of ammonium chloride compared to the mass of ammonium chloride. Draw a line of best fit for the points.