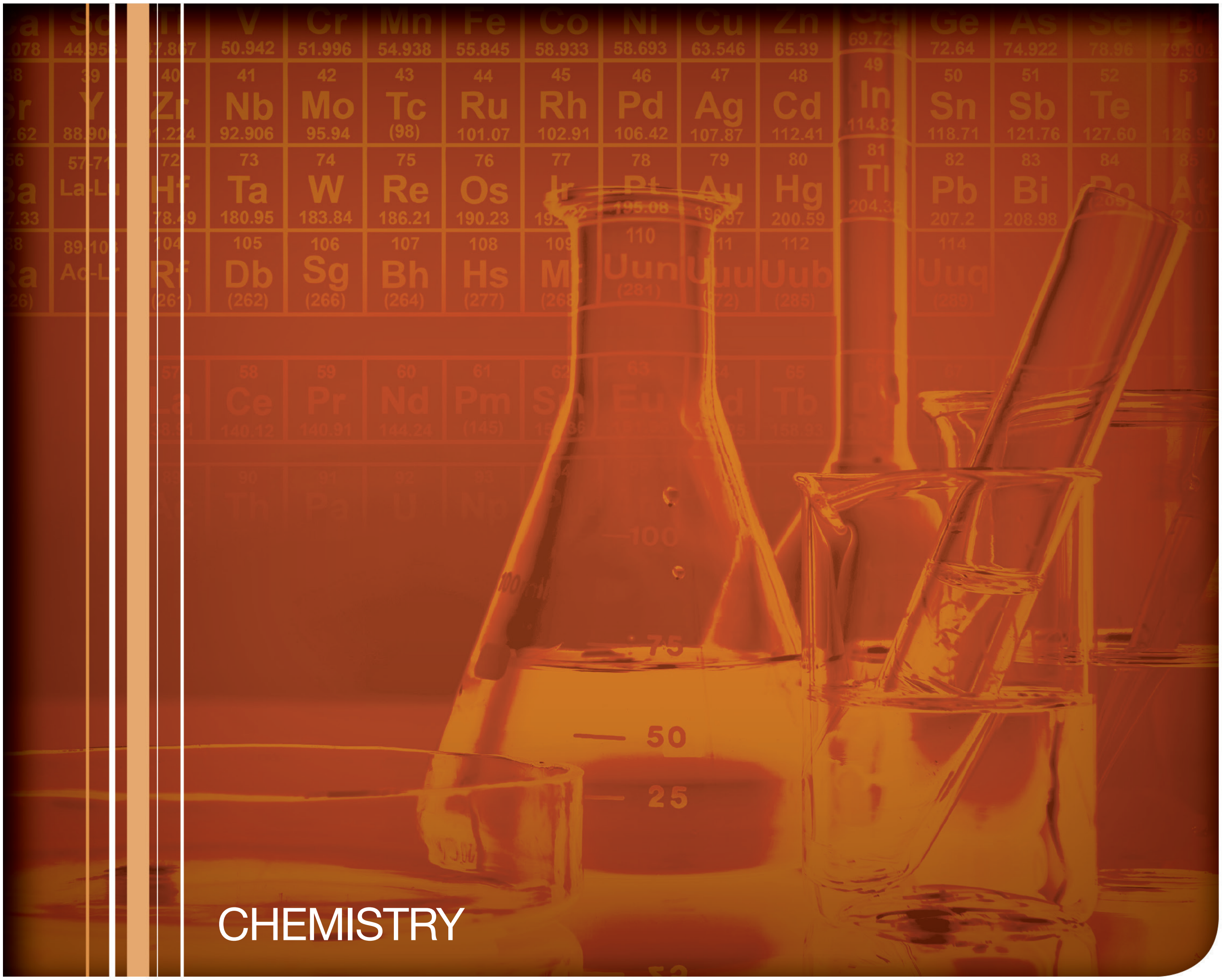
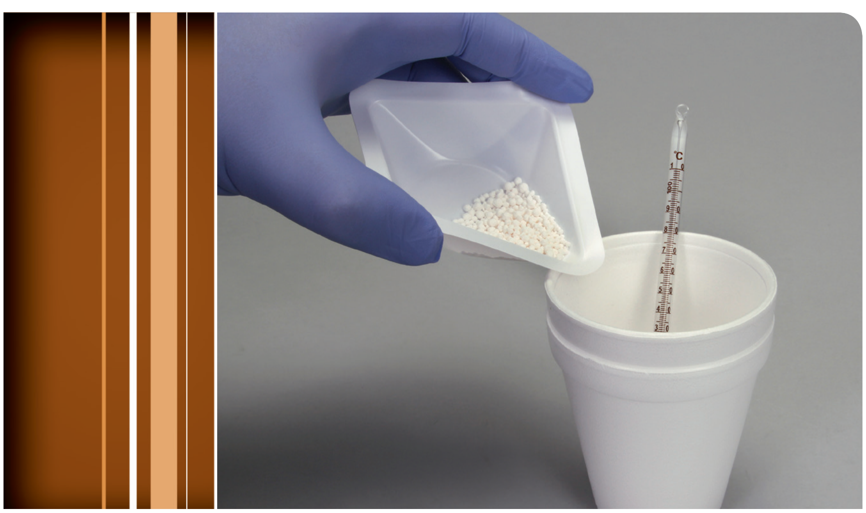


The Fundamentals of Calorimetry



Investigation Manual

CAROLINA
DISTANCE LEARNING

THE FUNDAMENTALS OF CALORIMETRY

Table of Contents

Overview

This investigation uses calorimetry to measure the heat of solution for two salt compounds commonly used in hot packs and cold packs. You will perform a graphical analysis of the calorimetry data for an endothermic and exothermic salt and predict how many grams of each salt are required to achieve a specific temperature when activated. After determining the heats of solution for both compounds, you will design a proposal for constructing a hot pack and a cold pack containing 100 g of water.

Outcomes

- Calculate the heat capacity of the calorimeter.
- Calculate the enthalpy of solution for ammonium chloride and calcium chloride.
- Design a chemical hot pack and cold pack that meets given volume and temperature specifications for first aid treatment.

Time Requirements

Preparation .... 15 minutes
Activity 1: Heat Capacity of the Calorimeter .... 30 minutes
Activity 2: Enthalpy of Solution for Calcium Chloride .... 30 minutes
Activity 3: Enthalpy of Solution for
Ammonium Chloride .... 30 minutes
Activity 4: Design a Proposal for a Hot Pack and a
Cold Pack .... 15 minutes

Key

Personal protective equipment (PPE)

goggles gloves apron

follow link to video

photograph results and submit

stopwatch required

warning corrosion flammable toxic environment health hazard

Made ADA compliant by NetCentric Technologies using the CommonLook® software

CAROLINA®

Background

When people sprain their ankle they are often given an instant cold pack to prevent swelling at the injury site. After 48–72 hours, a hot pack can be applied to relieve pain and stimulate blood flow for faster healing. These packs are activated by breaking an internal water pouch, which mixes with a salt contained in the outer bag. The amount of heat absorbed or released depends on the salt used and can be determined by the science of chemical thermodynamics.

Chemical thermodynamics is the study of energy changes that accompany physical and chemical transformations. These changes occur through the generation or absorption of a quantifiable amount of heat (q). If heat is generated during the reaction or physical change, it is an exothermic reaction. If the heat is absorbed or needs to be added to the reaction or physical change, then it is called an endothermic reaction.

Scientists use the term enthalpy (H) to measure the total energy of a system at constant pressure. Enthalpy also accounts for any change in volume during a reaction, which is important when dealing with gas reactions. The equation for the enthalpy change of a reaction is:

$$\Delta H = \Delta q + P\Delta V$$

Where ΔH is the change in enthalpy, Δq is the heat change, P is pressure (which remains constant), and ΔV is the change in volume. In most calorimeters, pressure and volume are constant. Therefore, the equation becomes:

$$\Delta H = \Delta q$$

Enthalpy is dependent on the amount of substance present. The enthalpy change for a reaction is generally written as a molar quantity. To calculate the molar enthalpy of a reaction, divide the reaction heat by the number of moles of reactant or product:

$$\Delta H_{rxn} = q_{rxn}/mol$$

Heat energy is always conserved within a system and all the parts of that system will adjust so that everything is the same temperature. For example, if a hot piece of copper metal is added to room temperature water, the heat energy transfers from the copper to the water. The energy transferred from the copper to the water because of the temperature difference is the heat (q). If the system is left undisturbed, the energy will continue to flow from the copper to the water until both are the same temperature, indicating that the system is at equilibrium.

Calorimetry is the act of measuring heat energy. The magnitude and the direction of heat transfer can be determined by using a calorimeter. A calorimeter is an insulated apparatus designed to create a closed system by preventing heat from flowing in or out. In this lab the calorimeter is two foam coffee cups stacked together and placed in a 250-mL beaker. The double foam walls and the air surrounding the cups in the beaker insulate the system.

In this investigation, the temperature change that occurs when a salt is dissolved in water will be measured. The heat released or absorbed during the reaction will be calculated by measuring the temperature change. However, heat is not the same as temperature. Temperature measures how hot or cold the sample is and is independent of mass. Heat or heat energy

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Background continued

measures the energy a sample contains. It is an extensive property and depends on the material quantity (mass). For example, a drop of boiling water and a gallon of boiling water both have a temperature of 100 °C, but the gallon of water has significantly higher heat energy because there is more of it.

The amount of heat that is given off or absorbed by a substance depends on both the mass and the substance composition. Every substance has its own **specific heat** (C_p), which is defined as the amount of heat energy necessary to raise the temperature of 1 gram by 1 °C. The specific heat of water is 4.18 J/g °C, which means it takes 4.18 joules of energy to raise the temperature of 1 g of water by 1 °C. More examples of specific heat include aluminum, which has a specific heat of 0.22 cal/g °C or 0.91 J/g °C, and copper whose specific heat is 0.09 cal/g °C or 0.385 J/g °C. Water requires more energy to heat than aluminum and copper metal.

In the calorimeter, a measured mass of water at a given temperature will be placed in the inner cup and a chemical will be added. When the chemical dissolves in the water the heat can either be transferred to the water or absorbed from the water by the act of chemical solvation. Since this reaction is carried out at constant (atmospheric) pressure, the heat of the reaction can be calculated using the change in temperature of the water.

The heat transfer or change in enthalpy for a reaction (q_{rxn}) is related to the mass of the solution (m), the specific heat capacity of the solution at constant atmospheric pressure (C_p), and the temperature change ($\Delta T = T_{final} - T_{initial}$) as shown in Equation 1. The minus sign on the right side of the equation means that if a solute releases heat, the water will absorb it, and if a solute absorbs heat, the water will release it. If a reaction absorbs heat, it is considered **endothermic**. This means that the reaction will be observed to cool. If a reaction releases heat it is **exothermic**, meaning that the temperature increased.

Equation 1.

$$q_{rxn} = -(m \times C_p \times \Delta T)$$

No calorimeter is a perfect insulator. There will be a small amount of energy transferred to and from the calorimeter. If the temperature of the reaction solution inside the calorimeter is higher than the calorimeter, some heat will be lost from the solution to heat up the calorimeter. Therefore, a calculation for the calorimeter itself must be added to the solution calculation on the right side of Equation 1, as shown in Equation 2.

Equation 2.

$$q_{rxn} = -[(m_{solution} \times C_{p\ solution} \times \Delta T_{solution}) + (m_{calorimeter} \times C_{p\ calorimeter} \times \Delta T_{calorimeter})]$$

Since the mass and specific heat of the calorimeter remain constant, they can be multiplied together to form another constant ($m \times C_p = C$). C represents the heat capacity of the calorimeter with the unit J/ °C in Equation 3.

Equation 3.

$$q_{rxn} = -[(m_{water} \times C_{p\ water} \times \Delta T_{water}) + (C_{calorimeter} \times \Delta T_{calorimeter})]$$

The change in temperature for the calorimeter is assumed to be the same as the change in temperature for water since they are in contact with each other.

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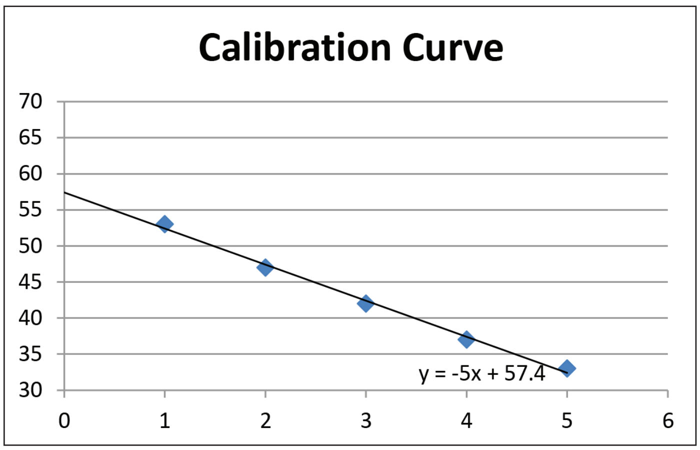
Sample Calculations

In Activity 1, the heat capacity of the calorimeter will be determined by analyzing the temperature change of the cold water in the calorimeter when hot water is added. The difference between the heat lost by the hot water and the heat gained by the cold water will be due to the heat absorbed by the calorimeter.

Figure 1.



A calibration curve is used to determine the heat capacity of the calorimeter. The calibration curve is a plot of the water temperature in the calorimeter as a function of time. Plotting the temperature as a function of time will determine the cooling rate of the mixture. You can accurately measure the initial mixing temperature by fitting the data with a linear trend line using Excel. Figure 1 shows a calibration curve generated by adding 50 g of 96 °C water to 50 g of 24 °C water. The initial temperature of the mixture should have been 60 °C [(96 + 24)/2 = 60]. The y-intercept from the linear trend curve is 57.4 °C and will represent T_{final}.

$$\Delta T_{(hot\ water)} = 57.4\ °C - 96\ °C = -38.6\ °C$$

$$\Delta T_{(cold\ water)} = 57.4\ °C - 24\ °C = -33.4\ °C$$

The energy lost by this reaction ($\Delta q_{(hot\ water)}$ + $\Delta q_{(cold\ water)}$) will be gained by the calorimeter ($\Delta q_{(cal)}$), **so the value of $\Delta q_{(cal)}$ should be set to be positive** in Figure 2.

Figure 2.

$\Delta q = C \times \Delta T \times m$
$\Delta q_{(hot\ water)} = 4.18$ J/g °C × –38.6 °C × 50 g = –8,067 J
$\Delta q_{(cold\ water)} = 4.18$ J/g °C × 33.4 °C × 50 g = 6,981 J
$\lvert \Delta q_{(hot\ water)} + \Delta q_{(cold\ water)} \rvert = \Delta q_{(cal)}$
Energy gained by calorimeter = 1,086 J
$\Delta q_{(cal)}$ = 1,086 J = $C_{(cal)}$ × 33.4 °C
C_{cal} = 32.5 J/°C

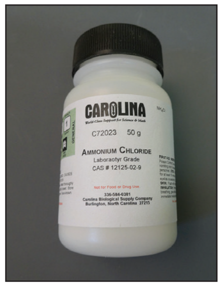
THE FUNDAMENTALS OF CALORIMETRY

Materials

Included in the materials kit:

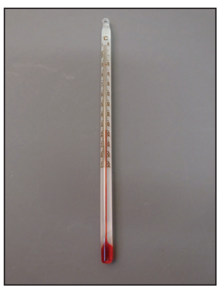
3 Foam cups, 8 oz

Calcium chloride, anhydrous, 50 g



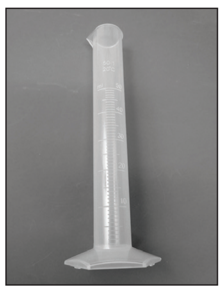
Ammonium chloride, 50 g

Needed from the equipment kit:



Thermometer

2 Weigh boats



Graduated cylinder, 50 mL

2 Plastic spoons

Electronic balance

Beaker, 250 mL

Needed, but not supplied:

- Water, bottled or purified
- Source of hot water (70–80 °C)
- Timing device
- Paper towels

There are many solids found in tap water that will affect your calculations. Bottled, purified, or filtered water from a home water purifier (e.g., Brita® or PUR®) should be used.

Reorder Information: Replacement supplies for The Fundamentals of Calorimetry investigation can be ordered from Carolina Biological Supply Company, kit 580321.

Call 800-334-5551 to order.

Safety

Goggles, gloves, and a lab apron are required at all times while conducting this investigation.

Read all the instructions for this laboratory activity before beginning. Follow the instructions closely, and observe established laboratory safety practices, including the use of appropriate personal protective equipment (PPE) as described in the Safety and Procedure sections.

Use caution when heating water to 75–80 °C in a microwave. Heat in short intervals, and check the temperature with a thermometer. Do not exceed 80 °C. Higher temperatures may melt the plastic graduated cylinder when measuring the water.

Calcium chloride and **ammonium chloride** are harmful if swallowed and can cause serious eye irritation.

Do not eat, drink, or chew gum while performing this activity. Wash your hands with soap and water before and after performing the activity. Clean up the work area with soap and water after completing the investigation. Keep pets and children away from lab materials and equipment.

Preparation

1. Read through the procedure.
2. Collect all the materials.
3. Select a clean work area.
4. Construct a calorimeter by placing one foam cup inside another and placing these two cups inside a 250-mL beaker. Placing the two foam cups inside the beaker prevents the calorimeter from tipping over.

ACTIVITY 1

A Heat Capacity of the Calorimeter

1. Measure 50.0 mL of water with a graduated cylinder and pour into the inner foam cup of the calorimeter. This will be the cold water.
2. Place the foam cup with the cold water inside the beaker.

> Remember that 1 mL of water = 1 g of water.

3. Stir the water with the thermometer until temperature is constant and record it as the initial temperature of the cold water (T_c) in Data Table 1.
4. Fill a third foam cup ¾ full with the hottest tap water possible.
5. Take the temperature of the hot water. If the temperature is less than 80 °C, heat in a microwave with short bursts until it reaches 80 °C.
6. Measure 50.0 mL of the hot water with a graduated cylinder. Place a thermometer in the cylinder and gently stir to get a constant reading. The temperature should be 75–80 °C. Record this as the temperature of the hot water (T_h) in Data Table 1.
7. Quickly pour the 50.0 mL of hot water into the cold water and start a timer.
8. Gently stir the mixed water with the thermometer, and record the temperature at 1 minute and then at 1-minute intervals for 10 minutes in Data Table 1.
9. Discard the water, dry the cup with a paper towel, and place it back in the calorimeter.
10. Repeat this activity, and collect a second set of data.

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Data Table 1: The Heat Capacity of the Calorimeter

Time (min)	Trial 1 Temp. °C	Trial 2 Temp. °C
1		
2		
3		
4		
5		
6		
7		
8		
9		

	Trial 1	Trial 2
Initial temperature of cold water, T_c		
Initial temperature of warm water, T_h		
Temperature at time 0 from graph, T_0		
Heat lost by hot water $\Delta q_{(hot\ water)} = C \times \Delta T \times m$		
Heat gained by cold water $\Delta q_{(coldwater)} = C \times \Delta T \times m$		
Heat gained by calorimeter in J $\Delta q_{(cal)} = \lvert \Delta q_{(hot\ water)} + \Delta q_{(cold\ water)} \rvert$		
Temperature change of the calorimeter $\Delta T_{cal} = T_0 - T_c$		
Heat capacity (C) of the calorimeter in J/°C $C_{cal} = \Delta q_{(cal)} / \Delta T_{cal}$		
Average heat capacity (C) of calorimeter in J/°C		

11. Create a spreadsheet and graph for the temperature versus time data.

12. Calculate the linear trend line and use the y-intercept to find the temperature at time 0 (T_0) when the two volumes of water are mixed. Record this value in Data Table 1.

13. Calculate the heat lost by hot water and record in Data Table 1.
$\Delta q_{(hot\ water)} = C \times \Delta T \times m$

14. Calculate the heat gained by cold water and record in Data Table 1. $\Delta q_{(coldwater)} = C \times \Delta T \times m$

15. Calculate the heat gained by the calorimeter and record in Data Table 1.
$\Delta q_{(cal)} = |\Delta q_{(hot\ water)} + \Delta q_{(cold\ water)}|$

16. Calculate the temperature change of the calorimeter and record in Data Table 1.
$\Delta T_{cal} = T_0 - T_c$

17. Calculate the heat capacity (C) of the calorimeter in J/°C and record in Data Table 1.
$C_{cal} = \Delta q_{(cal)} \ / \ \Delta T_{cal}$

18. Calculate the average heat capacity of the calorimeter and record in Data Tables 1 and 2.

ACTIVITY 2

A Enthalpy of Solution for Calcium Chloride

1. Measure 100 mL of pure or bottled water with a graduated cylinder, and pour it into the calorimeter.
2. Stir the water with a thermometer until the temperature is constant, and record it as the initial temperature (T_i) in Data Table 2.
3. Weigh approximately 5 g of calcium chloride ($CaCl_2$). Record the exact mass in Data Table 2.
4. Add the calcium chloride to the calorimeter, and stir the solution with the thermometer.
5. Keep stirring until the temperature stops increasing. Record this final temperature as T_f in Data Table 2.
6. Discard the solution, and rinse the inner cup. Dry the cup, and place back in the calorimeter.
7. Repeat steps 1–6 with 10.0 g and 15.0 g of $CaCl_2$.

ACTIVITY 3

A Enthalpy of Solution for Ammonium Chloride

1. Measure 100 mL of pure or bottled water with a graduated cylinder, and pour it into the calorimeter.
2. Stir the water with a thermometer until the temperature is constant, and record it as the initial temperature (T_i) in Data Table 2.
3. Weigh approximately 5 g of ammonium chloride (NH_4Cl). Record the exact mass in Data Table 2.
4. Add the ammonium chloride to the calorimeter, and stir the solution with the thermometer.
5. Keep stirring until the temperature stops decreasing. Record this final temperature as T_f in Data Table 2.
6. Discard the solution, and rinse the inner cup. Dry the cup, and place back in the calorimeter.

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ACTIVITY 3 continued

7. Repeat steps 1–6 with 10.0 g and 15.0 g of NH_4Cl.
8. For all trials in activities 2 and 3, calculate the following and record in Data Table 2:
9. Moles of Salts
   Moles = mass/molar mass
   molar mass $CaCl_2$ = 110.98 g/mol
   molar mass NH_4Cl = 53.49 g/mol
10. Change in temperature (°C)
    $\Delta T = T_f - T_i$
11. Heat absorbed by the solution (J)
    $c_w = 4.18$ (J/g)°C
    $q_w = -[c_w \times m_w \times \Delta T]$
12. Heat absorbed by the calorimeter (J)
    $q_{cal} = -[C_{cal} \times \Delta T]$
13. Enthalpy of solution (J)
    $\Delta H = q_w + q_{cal}$
14. Enthalpy of solution (kJ)
    ΔH (kJ) = ΔH/1000
15. Enthalpy/mole of solution (kJ/mol)
    $\Delta H_{sol} = \Delta H$ (kJ) / moles of salt
16. Average enthalpy/mole of solution (kJ/mol), you will have 1 average for your $CaCl_2$ trials and another for your NH_4Cl trials.

ACTIVITY 4

A Design a Proposal for a Hot Pack and a Cold Pack

Based on the data in Data Table 2 for calcium chloride and ammonium chloride, determine which compound to use and what quantity of each compound will be needed to make a chemical hot pack and cold pack. Both packs should be calculated based on using 100 g of water. The hot pack should reach 60 °C, and the cold pack should go down to 3.0 °C from a room temperature of 25 °C.

Create a spreadsheet and graph for $CaCl_2$ and NH_4Cl. Plot the mass on the x-axis and change in temperature on the y-axis for both graphs. The slope will be the change in temperature per gram of salt dissolved.

Hot Pack: Compound needed to achieve 60 °C: ____________________

Grams needed per 100 g of water: ________

Cold Pack: Compound needed to achieve 3.0 °C: ____________________

Grams needed per 100 g of water: ________

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Data Table 2: The Heat of Solution for Calcium Chloride and Ammonium Chloride

	Calcium Chloride			Ammonium Chloride		
	5 g $CaCl_2$	10 g $CaCl_2$	15 g $CaCl_2$	5 g NH_4Cl	10 g NH_4Cl	15 g NH_4Cl
Mass of water (g)						
Mass of salt (g)						
Moles of salt (g × mol/g)						
Initial temperature (°C) T_i						
Final temperature (°C) T_f						
Change in temperature (°C) $\Delta T = T_f - T_i$						
Heat absorbed by the solution (J) $q_w = -[c_w \times m_w \times \Delta T]$						
Heat capacity of the calorimeter (J/°C)						
Heat absorbed by the calorimeter (J) $q_c = -[C \times \Delta T]$						
Enthalpy of solution (J) $\Delta H = q_w + q_{cal}$						
Enthalpy of solution (kJ)						
Enthalpy/mole of solution (kJ/mol)						
Average enthalpy/mole of solution (kJ/mol)						

Disposal and Cleanup

1. Solutions of $CaCl_2$ and NH_4Cl can be poured down the drain with running water.
2. Discard all three foam cups and plastic spoons.
3. Rinse the beaker and dry with a paper towel.
4. Clean and sanitize the work space.

CHEMISTRY
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