

EXPERIMENT (1): THERMAL EQUILIBRIUM AND THE ZEROth LAW

Aim

The aim of this experiment is verification of the zeroth law of thermodynamics.

Introduction

The zeroth law of thermodynamics states that if both of two thermodynamic systems are in thermal equilibrium with a third, then both are in thermal equilibrium with each other. Accordingly, thermal equilibrium between systems is a transitive relation. Two systems are said to be in thermal equilibrium if they have the same temperature.

Materials

Saturated NaCl aqueous solution at lab temperature

Hot water

water

500 ml- beakers

Two 100 ml- beakers

Thermometer



Procedures

- 1) In a suitable container (large beaker or a large dish), put hot water to a high temperature (between 50 °C and 70 °C). Call this water “system (C)”.
- 2) In a 100 ml- beakers put about 50 ml of water at lab temperature. Call this “system (A)”.
- 3) In another 100 ml- beaker put about 50 ml the saturated NaCl solution at laboratory temperature. Call this “system (B)”.
- 4) Put both systems (A) and (B) inside the system (C).
- 5) After some minutes measure the temperatures of system (A) and system (C), and once both systems have the same temperature record the values as $T_{\text{system (A)}}$ and $T_{\text{system (C)}}$.
- 6) Also, measure the temperature of system (B) and system (C). Record the values as $T_{\text{system(A)}}$ and $T_{\text{system(C)}}$.

Results

Put your results in the following table:

	°C
$T_{\text{system (A)}}$	
$T_{\text{system (B)}}$	
$T_{\text{system (C)}}$	

Conclusion:

Write your conclusion in a way that verifies or does not verify the zeroth law of thermodynamics.

Questions

1. Define heat.
2. Can energy be stored in a system as heat? Rationalize.

REPORT OF EXPERIMENT (1)

Title:

Aim:

Results:

	°C
$T_{\text{system (A)}}$	
$T_{\text{system (B)}}$	
$T_{\text{system (C)}}$	

Conclusion:

Answers of the questions

EXPERIMENT (2): THE CONSTANT-PRESSURE CALORIMETRY

Aim

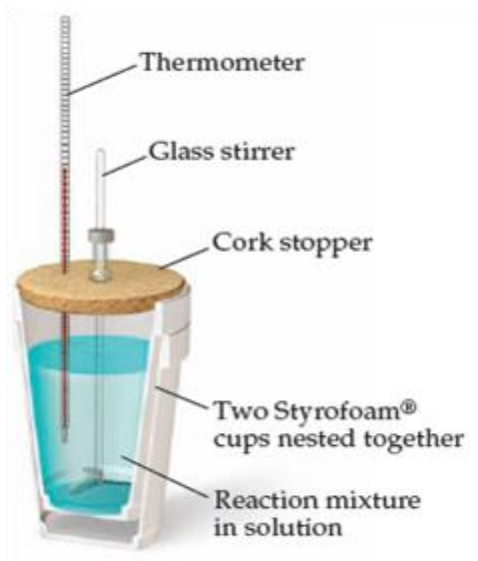
The aim of this experiment is to understand what a constant-pressure calorimetry means by the determination of the heat capacity of a constant-pressure calorimeter “ C_{cal} ”.

Introduction

- Calorimetry is a thermodynamic technique used to determine the heat transfer.
- The heat transfer is the quantity of heat absorbed or released by a system as a result of a physical or a chemical change in the thermodynamic state of the system.
- Physical and chemical changes are usually performed at constant temperature and also at either constant volume or constant pressure.
- If the change is performed at constant temperature and constant volume, the technique is called “constant-volume calorimetry”. It is also called “bomb calorimetry”. These kinds of changes with its calorimetry are rarely used and rarely preferred.
- If the change is performed at constant temperature constant pressure, the technique is called “constant-pressure calorimetry”, “coffee-cub calorimetry” or “Styrofoam calorimetry”. These kinds of calorimetry are the mostly used and the mostly preferred.
- In all of our experiments discussed and carried out here the technique that will be used is “The constant-pressure calorimetry”.
- The following figure shows the specially-manufactured constant-pressure calorimeter.



- The regular and common Styrofoam coffee cups are sometimes used to do the jobs the specially-manufactured constant-pressure calorimeter.



- The constant-Pressure calorimeter consists of two vessels (inner vessel and outer vessel) and two accessories. The inner vessel is accommodated by the outer vessel:
 - ✓ **The inner vessel (The calorimeter glass tube):**
The inner vessel is a glass tube with a size of about 50 mL. This glass has a mass and a known specific heat.
 - ✓ **The outer vessel:**
The outer vessel is an insulator (sometimes, it is called “jacket”) inside which the inner vessel is accommodated and kept protected from losing any of its energy as heat or intrusion of heat into it. This vessel usually is made of a Styrofoam that has walls with enough thickness so to ensure no transfer of heat across it. This outer vessel acts as an insulator of the inner vessel. The outer vessel also consists of a lid made of the same material and it has two holes to insert through the accessories.
 - ✓ **The accessories:**
The accessories are a thermometer and a stirrer both are inserted through the holes of the lid all the way to the bottom of the inner vessel.
- The system is put inside the inner vessel and is stirred gently to ensure that temperature is the same all over.
- The temperature of the system before anything happens must be accurately measured and recorded. This temperature is called the initial temperature.
- The system is then allowed to undergo its physical or its chemical change which will cause an increase or decrease in the temperature. During this change, gently stirring must continue in order to ensure that temperature is the same all over.
- The temperature of the system after the change is over must be accurately measured and recorded. This temperature is called the final temperature.
- In calorimetry, the equation used is the following equation:

$$q = C \times (T_2 - T_1)$$

$$q = m \times C_s \times (T_2 - T_1)$$

q = amount of heat gained by the system or lost (the unit is “J”)

C = the heat capacity of the system (the unit is “J/°C”)
 m = mass of the system (the unit is “g”)
 C_s = the specific heat of the system (the unit is “J/g °C”)
 t_1 = the initial temperature of the system (the unit is “°C”)
 t_2 = the final temperature of the system (the unit is “°C”)

- In this experiment a hot water will lose energy as heat and a cold calorimeter glass tube will gain this energy as heat.
- Lost heat q_{lost} will have a negative sign and gained heat q_{gained} will have a positive sign, but have the same value:

$$\begin{aligned}
 q_{\text{lost}} &= m_{\text{water}} \times 4.184 \text{ J/g } ^\circ\text{C} \times (\Delta t)_{\text{water}} \\
 q_{\text{gained}} &= C_{\text{cal}} \times (\Delta t)_{\text{cal}} \\
 C_{\text{cal}} \times (\Delta t)_{\text{cal}} &= - (m_{\text{water}} \times 4.184 \text{ J/g } ^\circ\text{C} \times (\Delta t)_{\text{water}}) \\
 C_{\text{cal}} &= - \frac{(m_{\text{water}} \times 4.184 \text{ J/g } ^\circ\text{C} \times (\Delta t)_{\text{water}})}{(\Delta t)_{\text{cal}}}
 \end{aligned}$$

Materials and chemicals

Constant pressure calorimeter with its accessories

100-mL graduated cylinder

Balance

Warm water (40 °C - 50 °C)

Procedure

- 1) Put the thermometer inside the empty inner vessel (The calorimeter glass tube) and measure the temperature. This temperature should be the same as the laboratory temperature. Record this temperature as “ t_{cal} ”.
- 2) Measure the mass of the inner vessel (The calorimeter glass tube). Record this mass as “ m_{cal} ”.
- 3) Measure the exact temperature of the hot water. Record this temperature as “ t_{water} ”.
- 4) Put inner vessel (The calorimeter glass tube) inside the outer vessel.
- 5) Up to no more than three quarter the space, fill the inner vessel (The calorimeter glass tube) with the hot water and immediately cover it with its lid.
- 6) Gently stirring, observe the thermometer and record it once it levels off. Record this temperature as “ t_{final} ”.
- 7) Measure the mass of the inner vessel (The calorimeter glass tube) with the water it contains. Record this mass as “ $m_{\text{cal+water}}$ ”.

Results

Tabulate your measurements as follows:

$m_{\text{cal+water}}/\text{g}$	m_{cal}/g	$t_{\text{cal}}/^\circ\text{C}$	$t_{\text{water}}/^\circ\text{C}$	$t_{\text{final}}/^\circ\text{C}$

Calculations

- 1) Calculate the mass of water:

$$m_{\text{water}} = m_{\text{cal+water}} - m_{\text{cal}}$$

- 2) Calculate the change in the temperature of the water:

$$(\Delta t)_{\text{water}} = t_{\text{final}} - t_{\text{water}}$$

- 3) Calculate the change in the temperature of the calorimeter glass tube:

$$(\Delta t)_{\text{cal}} = t_{\text{final}} - t_{\text{cal}}$$

- 4) Calculate the heat capacity of the calorimeter:

$$C_{\text{cal}} = \frac{(m_{\text{water}} \times 4.184 \text{ J/g } ^\circ\text{C} \times (\Delta t)_{\text{water}})}{(\Delta t)_{\text{cal}}}$$

- 5) Calculate the specific heat of the calorimeter, $C_{s,\text{cal}}$:

$$C_{s,\text{cal}} = \frac{C_{\text{cal}}}{m_{\text{cal}}}$$

Questions

- 1) Repeat your calculation but using for temperature the Kelvin unit instead of the Celsius unit. Do the final results differ from each other? Rationalize.
- 2) If the reference value of C_{cal} is $0.863 \text{ J/}^\circ\text{C g}$, calculate the experimental error.

REPORT OF EXPERIMENT (2)

Title:

Aim:

Results:

m_{cal}/g	$m_{\text{cal+water}}/\text{g}$	$t_{\text{cal}}/^{\circ}\text{C}$	$t_{\text{water}}/^{\circ}\text{C}$	$t_{\text{final}}/^{\circ}\text{C}$

Calculations

1) Mass of water:

2) Change in the temperature of the water:

3) Change in the temperature of the calorimeter glass tube:

4) Heat capacity of the calorimeter:

5) Specific heat of the calorimeter, $C_{s,\text{cal}}$:

Answers of the questions

1)

2) Calculation of the specific heat of the calorimeter, $C_{s,cal}$:

3) Calculation of the experimental error

EXPERIMENT 3: DETERMINATION OF SPECIFIC HEAT

Aim

The aim of this experiment is to determine the specific heat of a solid substance.

Introduction

- The heat absorbed or released, q , to change the temperature of a substance is given by the following equation:

$$q = m \times C_s \times (T_{\text{final}} - T_{\text{initial}})$$

“ m ” is the mass of a substance, “ C_s ” is its specific heat; “ T_{initial} ” is its initial and “ T_{final} ” is its final temperature.

- The heat capacity, C , of a substance is the quantity of heat required to raise the temperature of any mass “ m ” by only 1 °C:

$$C = m \times C_s$$

- Provided that no heat is exchanged with the surroundings, if a hot unknown metal becomes in contact with cold water, then heat will be lost from the metal and gained by the water and the calorimeter. The final temperature will be the same for all including the solid:

$$q_{\text{lost}} = -q_{\text{gained}}$$

- In this experiment a hot metal will be in contact with a cold water in a constant-pressure glass calorimeter. The following are the symbols that will be used in calculation:

Mass of the metal = m_{metal}

Specific heat of the metal = $C_{s,\text{metal}}$

Initial temperature the metal = T_{metal}

Mass water = m_w

Specific heat of water = $C_{s,w} = 4184 \text{ J/g } ^\circ\text{C}$

Mass of the glass calorimeter = m_{cal}

Specific heat of the glass calorimeter ($C_{s,\text{cal}}$) = $0.836 \text{ J/g } ^\circ\text{C}$

Initial temperature of water, calorimeter = $T_{w,\text{cal}}$

Final equilibrium temperature of mixture = T_{final}

Heat lost by the metal = q_{metal}

Heat gained by water, calorimeter and stirrer = $q_{w,\text{cal}}$

$$q_{\text{metal}} = m_{\text{metal}} \times C_{s,\text{metal}} \times (T_{\text{final}} - T_{\text{metal}})$$

$$q_{w,\text{cal}} = \{m_{\text{water}} \times C_{s,\text{water}} \times (T_{\text{final}} - T_{w,\text{cal}})\} + \{m_{\text{cal}} \times C_{s,\text{cal}} \times (T_{\text{final}} - T_{w,\text{cal}})\}$$

$$q_{w,\text{cal}} = \{(m_{\text{water}} \times C_{s,\text{water}}) + (m_{\text{cal}} \times C_{s,\text{cal}})\} \times (T_{\text{final}} - T_{w,\text{cal}})$$

From the following equality, the specific heat of the metal ($C_{s,\text{metal}}$) can be obtained:

$$q_{\text{lost}} = -q_{\text{gained}}$$

$$\{m_{\text{metal}} \times C_{s,\text{metal}} \times (T_{\text{final}} - T_{\text{metal}})\} = -\{(m_{\text{w}} \times C_{s,\text{w}} + m_{\text{cal}} \times C_{s,\text{cal}}) \times (T_{\text{final}} - T_{\text{w,cal}})\}$$

$$C_{s,\text{metal}} = \frac{-\{(m_{\text{w}} \times C_{s,\text{w}} + m_{\text{cal}} \times C_{s,\text{cal}}) \times (T_{\text{final}} - T_{\text{w,cal}})\}}{\{m_{\text{metal}} \times (T_{\text{final}} - T_{\text{metal}})\}}$$

Materials and chemicals

- 1) An unknown metal (20 g – 30 g)
- 2) Balance
- 3) Thermometer
- 4) Constant-pressure calorimeter
- 5) 500 mL-beaker
- 6) Hotplate or Bunsen burner
- 7) Ice

Procedure

- 1) Fill the 500 mL-beaker about half way with hot water.
- 2) Weigh the solid metal. Record this mass as m_{metal} .
- 3) Insert the metal into the hot water using a thread.
- 4) Weigh the calorimeter tube. Record this mass as m_{cal} .
- 5) Fill the calorimeter tube about half way with cold water and add some crushed ice.
- 6) Weigh the calorimeter tube containing and the cold water. Record this mass as $m_{\text{cal+w}}$.
- 7) Place the calorimeter tube inside its insulation jacket (do not forget to cover with its lid).
- 8) Insert thermometer inside the calorimeter tube and measure the temperature of the cold water (Be sure to stir the water first). Record this temperature as $T_{\text{cal,w}}$.
- 9) To measure the temperature of the heated metal, insert thermometer inside the 500 mL-beaker which contains the hot water and the metal. Record this temperature as T_{metal} .
- 10) Without splashing any water, quickly and carefully transfer the heated solid from the hot water to the calorimeter tube. (Be sure to cover the tube with its lid as soon as this transfer is done).
- 11) Observe the thermometer, stir the content gently and once the temperature becomes almost constant record it as T_{final} .
- 12) Repeat all of the previous steps for a different metal if your instructor advises you to.

Results

Quantity:	Metal 1	Metal 2
Mass of the metal m_{metal} (g)		
Initial temperature of the metal, T_{metal} ($^{\circ}\text{C}$)		
Specific heat of the glass calorimeter $C_{s,\text{cal}}$ (J/g $^{\circ}\text{C}$)	0.836	0.836
Mass of calorimeter m_{cal} (g)		
Mass of calorimeter and water $m_{\text{cal+w}}$ (g)		
Specific heat of cold water. $C_{s,\text{w}}$ (J/g $^{\circ}\text{C}$)	4.184	4.184
Initial temperature of cold water, $T_{\text{cal,w}}$ ($^{\circ}\text{C}$)		

Final equilibrium temperature T_{final} (°C)		
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Calculation

- 1) Calculate mass of water:

$$m_w = m_{\text{cal,w}} - m_{\text{cal}}$$

- 2) From the following equation calculate the specific heat of the metal ($C_{s,\text{metal}}$):

$$C_{s,\text{metal}} = \frac{-\{(m_w \times C_{s,w} + m_{\text{cal}} \times C_{s,\text{cal}}) \times (T_{\text{final}} - T_{w+\text{cal}})\}}{\{m_{\text{metal}} \times (T_{\text{final}} - T_{\text{metal}})\}}$$

Questions

Answer the following questions:

- 1) Why were you asked to add pieces of ice to the water? Hint: Think of phase change and temperature.
- 2) What are the sources of error and how can it be minimized?
- 3) Why should the water be stirred?
- 4) By what method of heat transfer was the metal heated?
- 5) What is the purpose of using a calorimeter in this experiment?
- 6) Why is it important to transfer the hot metal quickly into the cold water?

REPORT OF EXPERIMENT (3)

Title:

Aim:

Results:

Calculations

1) Calculation of mass of water

2) Calculation the specific heat of the metal

Answers of the questions

1)

2)

3)

4)

EXPERIMENT 4: ENTHALPY OF HYDRATION OF SOLID SALTS (SODIUM ACETATE OR COPPER SULPHATE)

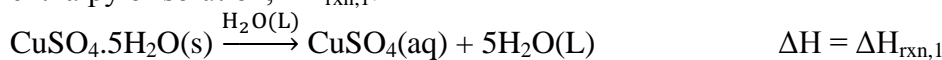
Aim

- The aim of this experiment is to measure the enthalpy of hydration of copper sulfate, sodium sulfate or any other salt that is capable to be hydrated.

Introduction

- The word “hydration” is used to mean different aspects. The process of dissolving a salt in water involves the attractions between the molecules of water and the ions of the salt. This is called “hydration”.
- When some salts crystallize they attach some water molecules to their solid crystal lattice. This process is also called “hydration”. This hydration is the subject of this experiment.
- Hydration is accompanied by heat exchange (absorption or releasing).
- Copper sulfate and sodium acetate are just examples of the salts which may have some water molecules within their solid crystalline lattice.
- The quantity of heat absorbed or released as a result of the attachment of water molecules to one mole of a solid ionic salt is called “the enthalpy of hydration”.
- The enthalpy of hydration is difficult to calculate as it is not possible to measure the chemical potential energy (the molar Gibbs energy) directly as the salt crystallizes. However, it can be determined using the following indirect method:

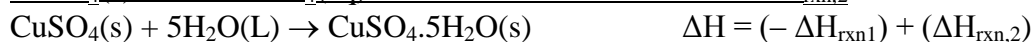
- ✓ First, we dissolve solid $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (or $\text{CH}_3\text{COONa} \cdot 3\text{H}_2\text{O}$) in water and evaluate its enthalpy of solution, $\Delta H_{\text{rxn},1}$:



- ✓ Second, we dissolve solid CuSO_4 in water and evaluate its enthalpy of solution, $\Delta H_{\text{rxn},2}$:



- ✓ Third, we apply Hess’s Law as follows:



Materials and chemicals

Scale

Constant-pressure calorimeter

Water

Sodium acetate (or Copper(II) sulfate)

Sodium acetate trihydrate (or copper(II) sulfate pentahydrate)

Procedure

Measure the mass of the calorimeter tube, Record this mass as “ m_{cal} ”.

Reaction 1

- In the calorimeter, add 50 mL of distilled water and measure the total mass. Record this mass as “ $m_{\text{cal}+\text{water},1}$ ”.

- Gently stir the water and measure the temperature once it settles off. Record this temperature as “ $T_{\text{initial},1}$ ”.
- Measure a mass between 5.00 g and 7.00 g of the hydrated salt (sodium acetate trihydrate, $\text{CH}_3\text{COONa}\cdot 3\text{H}_2\text{O}$ or copper sulfate pentahydrate, $\text{CuSO}_4\cdot 5\text{H}_2\text{O}$). Record the exact mass as “ $m_{\text{salt},1}$ ”.
- Add the hydrated salt to the water and cover the calorimeter with its lid. To dissolve the salt, stir the solution but gently.
- Observe the thermometer, and after being confident that all the mass is dissolved, measure the temperature as it settles off. Record this temperature as “ $T_{1,\text{final},1}$ ”.
- Record your experimental measurements on its correspondent table of results.

Reaction 2

Clean the calorimeter tube and dry it.

- As in “Reaction 1”, in the calorimeter, add 50 mL of distilled water and measure the total mass. Record this mass as “ $m_{\text{cal}+\text{water},2}$ ”.
- Gently stir the water and measure the temperature once it settles off. Record this temperature as “ $T_{\text{initial},2}$ ”.
- Measure a mass between 5.00 g and 7.00 g of the unhydrated salt (sodium acetate, CH_3COONa or copper sulfate, CuSO_4). Record the exact mass as “ $m_{\text{salt},2}$ ”.
- Add the hydrated salt to the water and seal the calorimeter with its lid. To dissolve the salt, stir the solution but gently.
- Observe the thermometer, and after being confident that all the mass is dissolved, measure the temperature as it settles off. Record this temperature as “ $T_{\text{final},2}$ ”.
- Record your experimental measurements on its correspondent table of results.

Results:

	Value
$m_{\text{cal}} / \text{g}$	

Reaction 1		Reaction 2	
	Value		Value
$m_{\text{cal}+\text{water},1} / \text{g}$		$m_{\text{cal}+\text{water},2} / \text{g}$	
$m_{\text{salt},1} / \text{g}$		$m_{\text{salt},2} / \text{g}$	
$m_{\text{solution},1} / \text{g}$		$m_{\text{solution},2} / \text{g}$	
$T_{\text{initial},1} / ^\circ\text{C}$		$T_{\text{initial},2} / ^\circ\text{C}$	
$T_{\text{final},1} / ^\circ\text{C}$		$T_{\text{final},2} / ^\circ\text{C}$	

Calculations:

(Note: Your teacher assistant will provide you with the values of the specific heats you need)

A) Reaction 1

- Calculate the mass of water:

$$m_{\text{water},1} = m_{\text{cal}+\text{water},1} - m_{\text{cal}}$$

2. Calculate the mass of solution:

$$m_{\text{solution},1} = m_{\text{water},1} + m_{\text{salt},1}$$

3. Calculate the temperature difference:

$$\Delta T_1 = T_{\text{final},1} - T_{\text{initial},1}$$

4. Calculate the heat exchanged:

$$q_1 = \{m_{\text{cal}} \times C_{s,\text{cal}} \times (\Delta T)_1\} + \{m_{\text{solution},1} \times C_{s,\text{solution},1} \times (\Delta T)_1\}$$

(Note that q_1 is the quantity of heat exchanged as a result of dissolving $n_{\text{salt},1}$ which is the number of moles of the hydrated salt that you dissolved.)

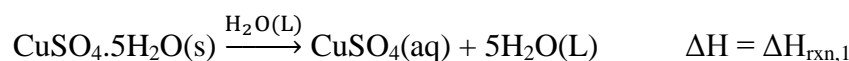
5. Calculate the number of moles of the salt:

$$n_{\text{salt},1} = \frac{m_{\text{salt},1}}{M_{\text{salt},1}}$$

6. Calculate, the change in enthalpy as a result of dissolving one mole of the salt:

$$\Delta H_{\text{rxn},1} = \frac{q_1}{n_{\text{salt},1}}$$

7. Write the thermochemical equation of dissolving the salt:



B) Reaction 2

1. Calculate the mass of water:

$$m_{\text{water},1} = m_{\text{cal+water},1} - m_{\text{cal}}$$

2. Calculate the mass of solution:

$$m_{\text{solution},2} = m_{\text{water},2} + m_{\text{salt},2}$$

3. Calculate the temperature difference:

$$\Delta T_2 = T_{\text{final},2} - T_{\text{initial},2}$$

4. Calculate the heat exchanged:

$$q_2 = \{m_{\text{cal}} \times C_{s,\text{cal}} \times (\Delta T)_2\} + \{m_{\text{solution},2} \times C_{s,\text{solution},2} \times (\Delta T)_2\}$$

(Note that q_2 is the quantity of heat exchanged as a result of dissolving $n_{\text{salt},2}$ which is the number of moles of the hydrated salt that you dissolved.)

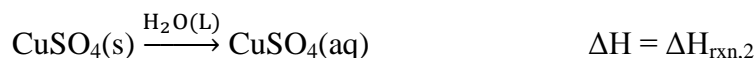
5. Calculate the number of moles of the salt:

$$n_{\text{salt},2} = \frac{m_{\text{salt},2}}{M_{\text{salt},2}}$$

6. Calculate, the change in enthalpy as a result of dissolving one mole of the salt:

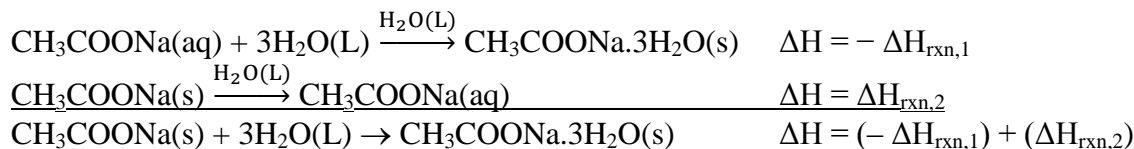
$$\Delta H_{\text{rxn},2} = \frac{q_2}{n_{\text{salt},2}}$$

7. Write the thermochemical equation of dissolving the salt:



C) Enthalpy of hydration of the salt

From the calculations of reaction 1 and reaction 2, apply Hess's law to calculate the enthalpy of hydration of the unhydrated salt:



D) Theoretical enthalpy of hydration of the salt

From the following literature information, calculate the theoretical enthalpy of hydration of solid sodium acetate, ΔH_{theor} :

	$\text{CH}_3\text{COONa} \cdot 3\text{H}_2\text{O}(\text{s})$	$\text{CH}_3\text{COONa}(\text{s})$	$\text{H}_2\text{O}(\text{L})$
$\Delta H_{\text{rxn}}^\circ / \text{kJ mol}^{-1}$	-1604.00	-709.32	-285.8

E) Error percentage

(Note: Your teacher assistant will provide you with the reference values of the enthalpy of hydration)

Calculate the error percentage of your experimental value of enthalpy of hydration:

$$\text{Error}\% = \pm \frac{\Delta H_{\text{theor}} - \Delta H_{\text{exp}}}{\Delta H_{\text{ref}}} \times 100$$

5. Calculation of the number of moles of the salt:

6. Calculation of, the change in enthalpy as a result of dissolving one mole of the salt:

7. The thermochemical equation of dissolving the salt:

B) Reaction 2

1. Calculate the mass of water:

2. Calculate the mass of solution:

3. Calculate the temperature difference:

4. Calculate the heat exchanged:

5. Calculate the number of moles of the salt:

6. Calculate, the change in enthalpy as a result of dissolving one mole of the salt:

7. Write the thermochemical equation of dissolving the salt:

C) Enthalpy of hydration of the salt

From the calculations of reaction 1 and reaction 2, apply Hess's law to calculate the enthalpy of hydration of the unhydrated salt:

D) Theoretical enthalpy of hydration

E) Error percentage

EXPERIMENT 5: THERMODYNAMICS OF THE OXIDATION OF ACETONE BY HYPOCHLORITE

Aim

The aim of this experiment is to measure the enthalpy of of the oxidation of acetone by hypochlorite.

Introduction

- ΔH is the heat that flows into or out of the system at constant temperature and pressure.
- By measuring the change in temperature and calculating how much reactant was consumed, we can determine *the enthalpy of reaction*.
- The reaction in this experiment is the oxidation of acetone by hypochlorite ion, the active species in bleach which will be used in this experiment:



- By performing the reaction in an insulated container open to the atmosphere and measuring the temperature change, the amount of heat absorbed or released can be determined. Thus, the enthalpy of reaction (ΔH_{rxn}) can be calculated.
- The following principles and formulas are used:
 - ✓ Any heat lost by the system is gained by the surroundings, and vice versa.
 - ✓ Heat absorbed or released (q) by a certain mass of the substance (m) can be calculate if we know the specific heat capacity of the substance (C_s) and the change in temperature (ΔT):

$$q = m \times C_s \times \Delta T$$

- The specific heat capacity of the calorimeter (the glass tube), $C_{s,\text{cal}}$ is known.
- The experiment involves the determination of the enthalpy of oxidation of acetone by hypochlorite (ΔH_{rxn}).

Apparatus and substances

- 1) Constant pressure calorimeter
- 2) Water bath (40 – 50) °C
- 3) 50-mL beaker
- 4) 250-mL beaker
- 5) Thermometer
- 6) 5-mL graduated pipette
- 7) 20-mL graduated pipette
- 8) 25 mL- graduated cylinder
- 9) 5.25% (w/v) sodium hypochlorite (Bleach)
- 10) Acetone aqueous solution (5% by volume acetone)

Procedure

(Warning: Bleach is irritating to the skin and dangerous if it gets in your eyes. Clean up any spills immediately).

- 1) Measure the mass of the calorimeter tub. Record this mass as m_{cal} .
- 2) Using the 20-mL pipette, add exactly 20.0 mL of 5.25% sodium hypochlorite solution (bleach) to the glass tube of the constant pressure calorimeter.
- 3) Measure the temperature of this solution. Record this temperature as T_{initial} .
- 4) Using the 5 mL-pipette, add 4.0 mL of 5.0% aqueous acetone into a 25-mL graduated cylinder.
- 5) Quickly open the calorimeter, pour the acetone solution into the bleach, and close the calorimeter.
- 6) Swirl the calorimeter gently while the reaction takes place (*do not use the thermometer as a stirring rod*).
- 7) Record the temperature as it levels off after *no less than* 10 minutes as T_{final} .
- 8) Repeat steps (1) – (6) two more times.

(Note: the quantities given in this procedure guaranties that acetone is the limiting reactants)

Results and calculations

- 1) Record your measurements in the following table

	1 st trial	2 nd trial	3 rd trial
$m_{\text{cal}} / \text{g}$			
$T_{\text{initial}} / ^\circ\text{C}$			
$T_{\text{final}} / ^\circ\text{C}$			

Calculations

- 1) Using the value of $C_{\text{s,cal}}$ you have already calculated in experiment 2, calculate the heat change:

$$q_{\text{cal}} = m_{\text{cal}} \times C_{\text{s,cal}} \times (T_{\text{final}} - T_{\text{initial}})$$

- 2) Because the concentration of acetone is 5% by volume acetone, it is clear that there is 5 mL of acetone in 100 mL of solution. Therefore, calculate the volume of acetone as follows:

$$\text{volume of acetone} = \frac{5}{100} \times \text{volume of solution}$$

- 3) From the volume of acetone, its density (0.791 g/mL) and its molar mass (58.078 g/mol), calculate the number of moles of acetone, n_{acetone} :

$$n_{\text{acetone}} = \frac{m_{\text{acetone}}}{M_{\text{acetone}}} = \frac{\text{volume of acetone} \times \text{density of acetone}}{M_{\text{acetone}}}$$

- 4) Knowing the amount of heat (q), and according to the following chemical equation:



Calculate ΔH_{rxn} in kJ / mole of acetone.

$$\Delta H_{\text{rxn}} = - \left(\frac{q_{\text{cal}}}{n_{\text{acetone}}} \right)$$

EXPERIMENT 6: ENTHALPY AND ENTROPY OF FUSION OF WATER

Aim

Understanding that fusion is a change of state of matter and its thermodynamic functions can be determined only if both states are present in equilibrium at constant temperature and pressure.

Introduction

The enthalpy of fusion

- Enthalpy of fusion is the heat absorbed by a substance to change the substance from its solid state to its liquid state at constant pressure and temperature.
- If the pressure is the standard pressure this change of state occurs at the substance standard melting point.
- Standard enthalpy of fusion, $\Delta H_{\text{fus}}^{\circ}$, is the heat absorbed by a substance to change its state from its solid state to its liquid state at standard pressure and at its melting temperature.
- This heat absorbed can be determined experimentally by calorimetry, and using the following equation:

$$q = m \times C_s \times (T_2 - T_1)$$

Where q , m , C_s , T_1 and T_2 are the heat absorbed, the mass, the specific heat, the initial temperature and the final temperature respectively.

The entropy of fusion

- Entropy of fusion is the change of the entropy of a substance as a result of changing it from its solid state to its liquid state at constant pressure and temperature.
- If the pressure is the standard pressure this change of state occurs at the substance standard melting point.
- Standard entropy of fusion is, $\Delta S_{\text{fus}}^{\circ}$, the change of the entropy of a substance as a result of changing it from its solid state to its liquid state at standard pressure and at its melting temperature.
- This change in the entropy can be determined, if the standard enthalpy of fusion is known, by the following equation:

$$\Delta S_{\text{fus}}^{\circ} = \frac{\Delta H_{\text{fus}}^{\circ}}{T_{\text{fus}}^{\circ}}$$

Substances and materials

Constant pressure calorimeter (glass tube, stirrer and Styrofoam container)

Hot-water bath (40 – 50) °C

Thermometer

Procedure

- Weigh the calorimeter glass tube. Record this mass as “ $m_{1,\text{tube}}$ ”.
- In a 250-mL beaker add distilled water up to about three quarters the beaker.

- Warm this water to a temperature that ranges between 40 °C and 60 °C.
- Fill the calorimeter glass tube up to about one third with the warm water.
- Again, weigh the calorimeter glass tube with the warm water. Record this mass as “m_{2,tube}”.
- Put the calorimeter glass tube with the warm water inside the Styrofoam. And insert the stirrer and the thermometer.
- Measure the temperature of the content of the calorimeter glass tube. Record this temperature as “T₁”.
- To the warm water inside the calorimeter glass tube add some pieces of ice.
- Stir the mixture gently.
- Make sure that the mixture contains water in both its liquid state and its solid state.
- Monitor the temperature of the content of the calorimeter glass tube, and once it levels off record the temperature as “T₂”.
- Quickly remove all ice from the calorimeter glass.
- Weigh the calorimeter glass tube with the liquid water inside. Record this mass as “m_{3,tube}”.

Results

$C_{s,\text{water}} = 4.184 \text{ J / g } ^\circ\text{C}$ $\Delta H_{\text{fus},\text{water}}^\circ = 6.01 \text{ kJ/mol}$ $\Delta S_{\text{fus},\text{water}}^\circ = 22 \text{ J/K}$ $T_{\text{fus},\text{water}}^\circ = 273 \text{ J/K}$ $M_{\text{water}} = 18.016 \text{ g/mol}$

- 1) Record your measurements in the following table

m _{1,tube} / g	m _{2,tube} / °C	m _{3,tube} / g	T ₁ / °C	T ₂ / °C

Calculations

- 1) Calculation the mass of the original hot water m_{original water}:

$$m_{\text{original water}} = m_{2,\text{tube}} - m_{1,\text{tube}}$$

- 2) Calculate the mass of the final water, m_{final water}:

$$m_{\text{final water}} = m_{3,\text{tube}} - m_{1,\text{tube}}$$

- 3) Calculate the mass of the melted ice, m_{melted ice}:

$$m_{\text{melted ice}} = m_{\text{final water}} - m_{\text{original water}}$$

- 4) Calculate the temperature difference, ΔT:

$$\Delta T = T_2 - T_1$$

- 5) From data given data and using the following equations calculate heat lost by the calorimeter tube and the hot water, q_{lost} :

$$q_{\text{lost}} = \{(m_{1,\text{tube}} C_{s,\text{tube}} (T_2 - T_1))\} + \{(m_{\text{origina water}} C_{s,\text{water}} (T_2 - T_1))\}$$

- 6) Calculate the heat gained by the ice q_{gained} :

$$q_{\text{gained}} = - (q_{\text{lost}})$$

- 7) From data given data, calculate number of moles of solid water (ice) melted, n_{ice} :

$$n_{\text{ice}} = \frac{m_{\text{melted ice}}}{M_{\text{water}}}$$

- 8) Calculate $\Delta H_{\text{fus,water}}^{\circ}$:

$$\Delta H_{\text{fus,water}}^{\circ} = \frac{q_{\text{gained}}}{n_{\text{ice}}}$$

- 9) From data given, calculate $\Delta S_{\text{fus,water}}^{\circ}$:

$$\Delta S_{\text{fus,water}}^{\circ} = \frac{\Delta H_{\text{fus,water}}^{\circ}}{T_{\text{fus,water}}^{\circ}}$$

- 10) From given data, calculate the percentage of your experimental error in the value of $\Delta S_{\text{fus,water}}^{\circ}$:

$$\text{error}\% = \frac{\Delta S_{\text{actual}}^{\circ} - \Delta S_{\text{exper}}^{\circ}}{\Delta S_{\text{actual}}^{\circ}} \times 100$$

REPORT OF EXPERIMENT (6)

Title:

Aim:

Results:

Calculations:

- 1) Calculation of the mass of the original hot water $m_{\text{original water}}$:

- 2) Calculation of the mass of the final water, $m_{\text{final water}}$:

- 3) Calculation of the mass of the melted ice, $m_{\text{melted ice}}$:

- 4) Calculation of the temperature difference, ΔT :

- 5) Calculation of the heat lost by the calorimeter tube and the hot water, q_{lost} :

6) Calculation of the heat gained by the ice q_{gained} :

7) Calculation of the number of moles of solid water (ice) melted, n_{ice} :

8) Calculation of $\Delta H_{\text{fus,water}}^{\circ}$:

9) Calculation of $\Delta S_{\text{fus,water}}^{\circ}$:

10) Calculation of the the percentage of your experimental error in the value of $\Delta S_{\text{fus,water}}^{\circ}$:

EXPERIMENT 7: FREE ENERGY AND TEMPERATURE EFFECT ON THE SOLUBILITY OF BORAX

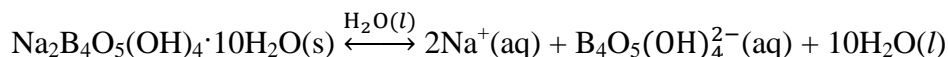
Title: Free energy and temperature effect on the solubility of borax

Aim

The aim of this experiment is the determination of thermodynamic parameters for making a saturated aqueous solution of borax.

Introduction

- The effect of temperature, the free energy, the enthalpy and the entropy of the solubility of borax in water will be determined.
- Borax, $\text{Na}_2\text{B}_4\text{O}_5(\text{OH})_4 \cdot 10\text{H}_2\text{O}$, is a sparingly soluble sodium salt in water. The equation of its solubility in water is presented by the following equation:



(Note: The two-heads arrow indicates that the solution is a heterogeneous solution which contains both the undissolved borax and the dissolved borax in a chemical equilibrium.)

- The equilibrium constant is presented as:

$$K = \frac{[\text{Na}^+(\text{aq})]^2 [\text{B}_4\text{O}_5(\text{aq})^{2-}] [\text{H}_2\text{O}(\text{l})]^{10}}{[\text{Na}_2\text{B}_4\text{O}_5 \cdot 10\text{H}_2\text{O}(\text{s})]}$$

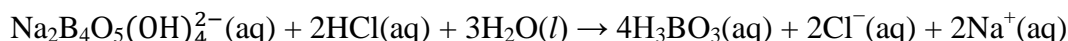
However, we know from our knowledge about heterogeneous equilibrium, that the concentration of any solid or liquid substance is constant. Therefore, the equation can be rearranged to take the following form:

$$K \times \frac{[\text{Na}_2\text{B}_4\text{O}_5 \cdot 10\text{H}_2\text{O}(\text{s})]}{[\text{H}_2\text{O}(\text{l})]^{10}} = [\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}] [\text{Na}^+(\text{aq})]^2$$

Because the left side of the last equation is a constant value, it is usually given a specific name and a specific symbol. The name is “the solubility product” and the symbol is “ K_{sp} ”. This convention turns the equilibrium constant equation for the solubility of this sparingly soluble salt to the following form:

$$K_{\text{sp, borax}} = [\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}] [\text{Na}^+(\text{aq})]^2$$

- The molarity of borate, $[\text{Na}_2\text{B}_4\text{O}_5(\text{OH})_4^{2-}(\text{aq})]$ can be calculate by titrating the borate ions with an aqueous solution of HCl that we know its molarity.
- The reaction between the borate and the aqueous HCl solution is as follows:



- The equation used to calculate the molarity of borate is as follows:

$$\left(\frac{M \times V}{n}\right)_{\text{borate}} = \left(\frac{M \times V}{n}\right)_{\text{HCl}}$$

$$M_{\text{borate}} = \left(\frac{M \times V}{n}\right)_{\text{HCl}} \times \left(\frac{n}{V}\right)_{\text{borate}}$$

- The chemical equation of borax dissolution given previously shows that the molarity of $\text{Na}^+(\text{aq})$ is twice the molarity of $\text{Na}_2\text{B}_4\text{O}_5(\text{OH})_4^{2-}(\text{aq})$:

$$[\text{Na}^+(\text{aq})] = 2 \times [\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}]$$

- Knowing the values of $[\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}]$ and $[\text{Na}^+(\text{aq})]$ enables us to calculate the value of the solubility product, $K_{\text{sp, borax}}$.
- In this experiment the value of $K_{\text{sp, borax}}$ will be calculated at two different temperatures (T_1 and T_2). This enables us to evaluate the following thermodynamic parameters:
 - We can calculate the standard change in the Gibbs free energy for dissolving borax in water at the two temperatures using the following equation:

$$\Delta G_{\text{solution}}^{\circ} = -R \times T \times \ln K_{\text{sp, borax}}$$

- We can calculate the standard change in the enthalpy of dissolving borax in water using the following equation:

$$\ln \frac{K_{\text{sp, borax, 1}}}{K_{\text{sp, borax, 2}}} = \frac{\Delta H_{\text{solution}}^{\circ}}{R} \times \left(\frac{T_1 - T_2}{T_1 \times T_2}\right)$$

(Note: The difference in the value of $\Delta H_{\text{solution}}^{\circ}$ at the two temperature is negligible.)

This equation can be rearranged to take the following form:

$$\Delta H_{\text{solution}}^{\circ} = \frac{\ln \frac{K_{\text{sp, borax, 1}}}{K_{\text{sp, borax, 2}}} \times R \times T_1 \times T_2}{T_1 - T_2}$$

- We can calculate the standard change in the entropy of dissolving borax in water at the two temperatures using the following equation:

$$\Delta G_{\text{solution}}^{\circ} = \Delta H_{\text{solution}}^{\circ} - T \times \Delta S_{\text{solution}}^{\circ}$$

This equation can be rearranged to take the following form:

$$\Delta S_{\text{solution}}^{\circ} = \frac{\Delta H_{\text{solution}}^{\circ} - \Delta G_{\text{solution}}^{\circ}}{T}$$

Materials and chemicals

Large and small beakers

Erlenmeyer flasks

Magnetic stirrer

Thermometers

A burette filled with standardized HCl solution (about 0.2 mol/L)

Procedure

Note

The experiment will may be carried out twice at two temperature:

The first at the lab temperature (about 25 °C)

The second is to be done at the lower temperature that equals to 5 °C or close to 5 °C

As the teacher assistant instructs, the experiment will may be carried out by a group.

- 1) In a two small beaker add about 20 g of the solid borax and about 400 mL distilled water.
(Note: The mass of the borax and the volume of the water are approximate and not necessarily exact but only to ensure the heterogeneity of the solution.)
- 2) Stir the mixture in for at least 10 minutes.
- 3) Place a thermometer in each beaker.
- 4) Put the first beaker on the bench at the lab temperature, and the second in a larger beaker that contains water and some ice.
- 5) To allow the undissolved borax to settle to the bottom, leave both beakers undisturbed until the aqueous solution becomes clear.
- 6) Measure the temperature of the mixture in in each beaker. Record this temperature as “T₁” for the flask at the high temperature and as “T₂” for the flask at the low temperature.
- 7) Without disturbing the solid at the bottom, carefully pipet 10 mL of the solution in each beaker into an Erlenmeyer flask that is clean but not necessarily dry.
- 8) Add 20 mL of distilled water and 3 drops of methyl orange indicator to each of these three Erlenmeyer flasks. The solution of each should turn blue.
- 9) Titrate the sample with the standardize HCl solution. Record the volume of HCl at T₁ as V_{HCl,1} and at T₂ as V_{HCl,2}.
- 10) For each temperature, repeat the titration two more times.
- 11) Record the molarity of the standardize HCl solution as M_{HCl}.

Results

- 1) Temperatures and the molarity of the standardize HCl solution:

T ₁ / K	T ₂ / K	M _{HCl} / mol L ⁻¹

- 2) Titrations:

Titration number	V _{HCl,1} / mL	V _{HCl,2} / mL
1		
2		
3		
Average volume		

Calculations

1. Using the following equation calculate the molarity of the aqueous borate at T₁, [Na₂B₄O₅(aq)²⁻]₁ and at T₂, [Na₂B₄O₅(aq)²⁻]₂:

$$[\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}] = \left(\frac{M \times V}{n}\right)_{\text{HCl}} \times \left(\frac{n}{V}\right)_{\text{borate}}$$

2. Using the following equation calculate the molarity of the aqueous sodium ion at T₁, [Na(aq)⁺]₁ and at T₂, [Na(aq)⁺]₂:

$$[\text{Na}^+(\text{aq})] = 2 \times [\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}]$$

3. Using the following equation calculate the value of the solubility product at T₁, K_{sp,borax,1}, and at T₂, K_{sp,borax,2},

$$K_{\text{sp,borax}} = [\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}][\text{Na}^+(\text{aq})]^2$$

4. Also, using the following equation calculate the value of ΔG[°]_{solution} at T₁, ΔG[°]_{solution,1} and at T₂, ΔG[°]_{solution,2}:

$$\Delta G_{\text{solution}}^{\circ} = -R \times T \times \ln K_{\text{sp,borax}}$$

5. Finally, using the following equation calculate the value of ΔH[°]_{solution}:

$$\Delta H_{\text{solution}}^{\circ} = \frac{\ln \frac{K_{\text{sp,borax,1}}}{K_{\text{sp,borax,2}}} \times R \times T_1 \times T_2}{T_1 - T_2}$$

6. Using the following equation calculate the value of ΔS[°]_{solution} at T₁, and at T₂:

$$\Delta S_{\text{solution}}^{\circ} = \frac{\Delta H_{\text{solution}}^{\circ} - \Delta G_{\text{solution}}^{\circ}}{T}$$

Tabulate the results of these calculations as follows:

	At "T ₁ "	At "T ₂ "
[Na ₂ B ₄ O ₅ (aq) ²⁻]/mol L ⁻¹		
[Na ⁺ (aq)]/mol L ⁻¹		
K _{sp,borax}		
ΔG [°] _{solution} /kJ mol ⁻¹		
ΔH [°] _{solution}		
ΔS [°] _{solution} /J K ⁻¹ mol ⁻¹		

7. Using the following equation calculate the value of ΔS[°]_{solution} at T₁, and at T₂:

$$\Delta S_{\text{solution}}^{\circ} = \frac{\Delta H_{\text{solution}}^{\circ} - \Delta G_{\text{solution}}^{\circ}}{T}$$

REPORT OF EXPERIMENT (7)

Title

Aim:

Results:

1. Temperatures and the molarity of the standardize HCl solution:

2. Titrations:

Calculations

1. Calculation of $[\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}]_1$ and $[\text{Na}_2\text{B}_4\text{O}_5(\text{aq})^{2-}]_2$:

2. Calculation of $[\text{Na}(\text{aq})^+]_1$ and $[\text{Na}(\text{aq})^+]_2$:

3. Calculation of the solubility product, $K_{sp, borax,1}$ and $K_{sp, borax,2}$:

4. Calculation of $\Delta G_{solution}^{\circ}$ at T_1 , and at T_2 :

5. Calculation of $\Delta H_{solution}^{\circ}$:

6. Calculation the of $\Delta S_{solution}^{\circ}$ at T_1 , and at T_2 :

Table of the results of calculations:

EXPERIMENT 8: DETERMINATION OF THE ENTHALPY OF SOLUTION AT INFINITE DILUTION

Aim

Understanding the difference between the enthalpy of solution and the enthalpy of solution at infinite dilution " $\Delta H_{\text{infin.dilution}}^{\circ}$ ", and determining the value of $\Delta H_{\text{infin.dilution}}^{\circ}$.

Introduction

- Mole fraction, solute mass percent, molarity and molality are the most common ways of expressing the amount of a solute dissolved in a solvent. i.e. concentration.
- Dissolving a solid solute in a liquid solvent is usually accompanied by heat transfer between the solution and its surroundings. This heat transfer is called the enthalpy of solution, ΔH_{soln} .
- If the solution process is done at the standard pressure and the amount of solute is one mole, the heat transferred is called the standard enthalpy of solution, $\Delta H_{\text{soln}}^{\circ}$.
- The enthalpy of solution takes one of the following possibilities:
 - ✓ If the solution process is not accompanied by heat transfer from or to the system then ***the enthalpy of solution is zero*** and the solution is described as an ***ideal solution***. Such kind of solution is few and rare.
 - ✓ If the solution process leads the system to lose heat then the process is exothermic and ***the enthalpy of solution is negative*** and the solution is not an ideal solution.
 - ✓ If the solution process leads the system to absorb heat then the process is endothermic and ***the enthalpy of solution is positive*** and the solution, also, is not an ideal solution.
- The question which cannot be overlooked is the following question:
“if we dissolve 1 mole of a solid in a liquid, would the value of $\Delta H_{\text{soln}}^{\circ}$ be the same regardless of the quantity of the solvent? Or would it differ as the quantity of the solvent differs?”
- It is found that $\Delta H_{\text{soln}}^{\circ}$ differs as the quantity of the solvent differs. This is due to what is called “enthalpy of dilution.
- This means that the values of $\Delta H_{\text{soln}}^{\circ}$ cannot be tabulated easily. Therefore, chemist has to arrive to a solution to this case, and they did.
- Chemists take the value of the standard enthalpy of solution when the one mole is dissolved in a quantity of the solvent that is large enough so that beyond which no more heat transfer can take place.
- This quantity of the solvent differs as solute differs, but, in all cases, it is given a special name which is “the standard enthalpy of solution at infinite dilution”, and is given a characteristic symbol, namely “ $\Delta H_{\text{infin.dilution}}^{\circ}$ ”.
- As mentioned above, the aim of this experiment is to determine “ $\Delta H_{\text{infin.dilution}}^{\circ}$ ”.

Materials and chemicals

Balance

Constant-pressure calorimeter

A pipette, a burette or a graduated cylinder

$\text{NH}_4\text{Cl}(\text{s})$

Distilled water

Procedure

- 1) Measure the mass of the glass tube of the constant-pressure calorimeter. Record this mass as “ $m_{\text{glass tube}}$ ”.
- 2) Use a burette, pipette or graduated cylinder to add 50 mL distilled water into the glass tube. Record this volume as “ V_{water} ”.
- 3) Read the temperature and record it as “ T_1 ”.
- 4) Weigh about 5 g of the solute (NH_4Cl). Record this mass as “ m_{solute} ”.
- 5) In this step:
 First- *The solute you weighted must be added to the 50 mL in the glass tube. Note that once you add the solid, you must be sure that the solvent is completely dissolved. You may need to use the stirrer but gently.*
 Second- *The thermometer must be monitored and the temperature must read once it settles off. Record this temperature as “ T_2 ”.*
- 6) Clean the glass tube, the stirrer and the thermometer.
- 7) Repeat the steps from 2 to 6 two more times, using about 10 g and about 15 g of the solid ammonium chloride. Record these two masses as “ $m_{\text{solute},2}$ ” and “ $m_{\text{solute},3}$ ”.

Results

Density of water, $d_{\text{water}} = 1.0 \text{ g/mL}$
 Specific heat of solution, $C_{s,\text{soln}} = \text{specific heat of water, } C_{s,\text{water}} = 4.184 \text{ L/g } ^\circ\text{C}$
 Heat capacity of the calorimeter, C_{calor} . Will be given by your teacher assistant
 Molar mass of NH_4Cl , $M_{\text{solute}} = 53.489 \text{ g/mol}$

Tabulate the results of your experiment and calculations as follows:

	$m_{\text{cal}} / \text{g}$	$V_{\text{water}} / \text{mL}$	$m_{\text{salt}} / \text{g}$	$t_1 / ^\circ\text{C}$	$t_2 / ^\circ\text{C}$
1					
2					
3					

Calculations

Fill the following table after doing the calculations:

	$m_{\text{soln}} / \text{g}$	$\Delta t / ^\circ\text{C}$	q / J	$n_{\text{solute}} / \text{mol}$	$\Delta H / \text{kJ mol}^{-1}$	Molarity/ mol L^{-1}
1						
2						
3						

- 1) Calculate the mass of the water m_{water} :

$$m_{\text{water}} = V_{\text{water}} \times d_{\text{water}}$$

- 2) Calculate the mass of the solution, m_{soln} , used in the three trials:

$$m_{\text{soln}} = m_{\text{water}} + m_{\text{solute}}$$

- 3) Calculate the difference in temperature, Δt , in the three trials:

$$\Delta t = t_2 - t_1$$

- 4) Using the following equation, calculate the heat absorbed by the calorimeter and by the solution, q_{gained} in the three trials:

$$q_{\text{gained}} = (m_{\text{cal}} C_{s,\text{cal}} + m_{\text{soln}} \times C_{s,\text{soln}}) \times \Delta t$$

- 5) Calculate the number of moles of solute, n_{solute} , in the three trials:

$$n_{\text{solute}} = \frac{m_{\text{solute}}}{M_{\text{solute}}}$$

- 6) For each of the three trials, use the following equation to calculate the heat released per mole of the solute, $\Delta H_{\text{soln}}^{\circ}$:

$$\Delta H_{\text{soln}}^{\circ} = \frac{-q_{\text{gained}}}{n_{\text{solute}}}$$

- 7) Assuming that the volume of each solution, V_{soln} (L), equals the volume of water, V_{water} , and using the following equation, calculate the molarity, M , of each solution:

$$M_{\text{soln}} = \frac{n_{\text{solute}}}{V_{\text{soln}} \text{ (L)}}$$

- 8) Using a graph paper draw the relation between $\Delta H_{\text{soln}}^{\circ}$ on the Y axis and M_{soln} on the X axis.
9) Extrapolate the line of your graph to the point where it crosses the Y axis, and determine from this point the value of $\Delta H_{\text{infin.dilution}}^{\circ}$:

$$\Delta H_{\text{infin.dilution}}^{\circ} = \boxed{}$$

REPORT OF EXPERIMENT (8)

Title:

Aim

Results

Calculations

1) Calculation of m_{water} :

2) Calculation of m_{soln} :

3) Calculation of Δt :

4) Calculation of q_{gained} :

5) Calculation of n_{solute} :

6) Calculation of $\Delta H_{\text{soln}}^{\circ}$:

7) Calculation of M assuming that the volume of each solution, V_{soln} (L), equals the volume of water, $V_{\text{water}} = 0.05$ L:

Table of results of calculations:

8) From the attached graph and as extrapolate shows, the enthalpy of solution at infinite dilution is as indicated below:

$$\Delta H_{\text{infin.dilution}}^{\circ} = \boxed{}$$