جامعة الملك سعود
King Saud University
1957

COLLEGE OF SCIENCES CHEMISTRY
DEPARTMENT

CHEM 232

LABORATORY OF
CHEMICAL THERMODYNAMICS

DR. AHMAD AL-OWAIS

1446- 2024

TABLE OF CONTENTS

EXPERIMENT (1): THERMAL EQUILIBRIUM AND THE ZEROTH LAW

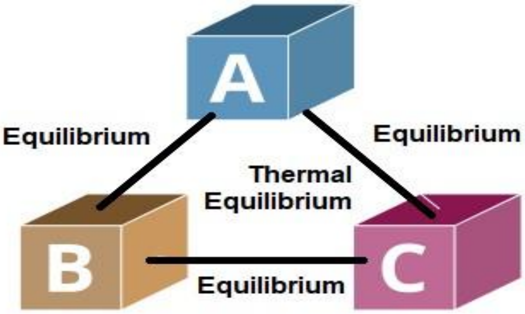
Aim

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.

The figure below shows that if system "A" is equilibrium with system "C " and system" B" also in thermal equilibrium with "C" , then the two systems "A" and "B" are in thermal equilibrium with eath other .



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 $T_{\text{system(C)}}$

REPORT OF EXPERIMENT (1)

Title:

Aim:

Results:

	°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.

Answers of the questions

1. Define heat.

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

EXPERIMENT (2): THE CONSTANT-PRESSURE CALORIMETRY

Aim

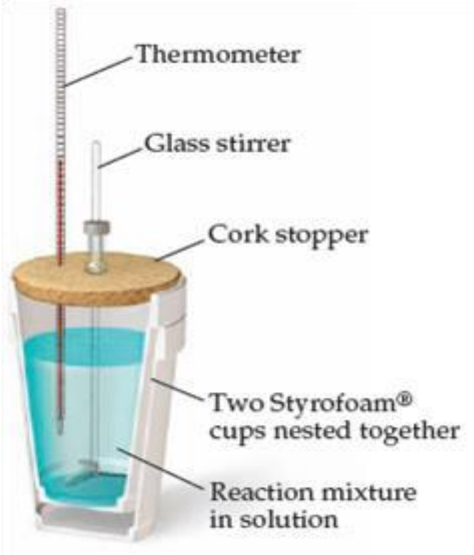
The aim of this experiment is to understand what a constant-pressure calorimetry means by the determination of the heat capacity of a constant-pressure calorimeter “C_{cal}”.

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:

$$q_{lost} = m_{water} \times 4.184\ J/g\ °C \times (\Delta t)_{water}$$

$$q_{gained} = C_{cal} \times (\Delta t)_{cal}$$

$$C_{cal} \times (\Delta t)_{cal} = -\ (m_{water} \times 4.184\ J/g\ °C \times (\Delta t)_{water})$$

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

Materials and chemicals

Constant pressure calorimeter with its accessories
graduated cylinder
Balance
Hot 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) 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 from water bath 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 "$m_{cal+water}$".

REPORT OF EXPERIMENT (2)

Title:

Aim:

Results:

m_{cal}/g	$m_{cal+water}$/g	t_{cal}/°C	t_{water}/°C	t_{final}/°C

Calculations

1) Calculate the mass of water:

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

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

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

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

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

4) Calculate the heat capacity of the calorimeter:

$$C_{cal} = -\frac{(m_{water} \times 4.184\ \mathrm{J/g\,°C} \times (\Delta t)_{water})}{(\Delta t)_{cal}}$$

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

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

6) Calculation of the experimental error If the reference value of $C_{s,cal}$ is 0.836 J/g.°C

$$\text{Error\%} = \left| \frac{C_{s,actual} - C_{s,exp}}{C_{s,actual}} \right| \times 100$$

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 $C_{s,cal}$ is 0.836 J/°C g, calculate the experimental error.

Answers of the questions

1) Repeating calculation using the Kelvin unit:

$(\Delta t)_{water}$ =

$(\Delta t)_{cal}$ =

C_{cal} =

$C_{s,cal}$: =

2) Error % in $C_{s,cal}$ =

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:

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

"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 ^{o}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:

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

- 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 = $C_{s,metal}$
Initial temperature the metal = T_{metal}
Mass water = m_w
Specific heat of water = $C_{s,w}$ = 4.184 J/g ^{o}C
Mass of the glass calorimeter = m_{cal}
Specific heat of the glass calorimeter ($C_{s,cal}$) = 0.836 J/g ^{o}C
Initial temperature of water, calorimeter = $T_{w,cal}$
Final equilibrium temperature of mixture = T_{final}
Heat lost by the metal = q_{metal}
Heat gained $q_{w,cal} = \{(m_{water} \times C_{s,water}) + (m_{cal} \times C_{s,cal})\} \times (T_{final} - T_{w,cal})$
by water, calorimeter and stirrer = $q_{w,cal}$

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

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

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

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

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

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

Materials and chemicals

An unknown metal
Balance
Constant-pressure calorimeter
500 mL-beaker
Hotplate or water bath
Thermometer
Ice

Procedure

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

Calculation

1) Calculate mass of water:

$$m_{w} = m_{cal,w} - m_{cal}$$

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

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

REPORT OF EXPERIMENT (3)

Title:

Aim:

Results:

Quantity:	Values
Mass of the metal m_{metal} (g)	
Initial temperature of the metal, T_{metal} (^{o}C)	
Specific heat of the glass calorimeter $C_{s,cal}$ (J/g °C)	0.836
Mass of calorimeter m_{cal} (g)	
Mass of calorimeter and water m_{cal+w} (g)	
Specific heat of cold water. $C_{s,w}$ (J/g °C)	4.184
Initial temperature of cold water, $T_{cal,w}$ (^{o}C)	
Final equilibrium temperature T_{final} (^{o}C)	

Calculations :

1) Calculation of mass of water :

2) Calculation the specific heat of the metal :

Questions :

1) What are the sources of error and how can it be minimized?

2) Why should the water be stirred?

3) What is the purpose of using a calorimeter in this experiment?

4) Why is it important to transfer the hot metal quickly into the cold water?

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 is in water involves the attractions between the molecules of water and the ions of the salt. This is called “hydration”.
- When some slats 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 $CH_3COONa.3H_2O$ in water and evaluate its enthalpy of solution, $\Delta H_{rxn,1}$:

  $CH_3COONa.3H_2O(s) \xrightarrow{H_2O(L)} CH_3COONa(aq) + 3H_2O(L) \qquad \Delta H = \Delta H_{rxn,1}$
- ✓ Second, we dissolve solid CH_3COONa in water and evaluate its enthalpy of solution, $\Delta H_{rxn,2}$:

  $CH_3COONa(s) \xrightarrow{H_2O(L)} CH_3COONa(aq) \qquad \Delta H = \Delta H_{rxn,2}$
- ✓ Third, we apply Hess’s Law as follows:

  $CH_3COONa(aq) + 3H_2O(L) \xrightarrow{H_2O(L)} CH_3COONa.3H_2O(s) \qquad \Delta H = -\Delta H_{rxn,1}$

  $\underline{CH_3COONa(s) \xrightarrow{H_2O(L)} CH_3COONa(aq) \qquad \Delta H = \Delta H_{rxn,2}}$

  $CH_3COONa(s) + 3H_2O(L) \rightarrow CH_3COONa.3H_2O(s) \qquad \Delta H = (-\Delta H_{rxn,1}) + (\Delta H_{rxn,2})$

Materials and chemicals

Constant-pressure calorimeter
Sodium acetate (or Copper(II) sulfate)
Sodium acetate trihydrate (or copper(II) sulfate pentahydrate)
Water
Scale (thermometer)

Procedure

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

Reaction 1

1. In the calorimeter, add 25 mL of distilled water and measure the total mass. Record this mass as “$m_{cal+water,1}$”.
2. Gently stir the water and measure the temperature once it settles off. Record this temperature as “$T_{initial,1}$”.
3. Measure a 2.5 g of the hydrated salt (sodium acetate trihydrate, $CH_3COONa.3H_2O$ or copper sulfate pentahydrate, $CuSO_4.5H_2O$). Record the exact mass as “$m_{salt,1}$”.
4. Add the hydrated salt to the water and cover the calorimeter with its lid. To dissolve the salt, stir the solution but gently.
5. Observe the thermometer, and after being confident that all the mass is dissolved, measure the temperature as it settles off. Record this temperature as “$T_{1,final,1}$”.
6. Record your experimental measurements on its correspondent table of results.

Reaction 2

Clean the calorimeter tube and dry it.

1. As in “Reaction 1”, in the calorimeter, add 25 mL of distilled water and measure the total mass. Record this mass as “$m_{cal+water,2}$”.
2. Gently stir the water and measure the temperature once it settles off. Record this temperature as “$T_{initial,2}$”.
3. Measure a 2.5 g of the unhydrated salt (sodium acetate, CH_3COONa or copper sulfate, $CuSO_4$). Record the exact mass as “$m_{salt,2}$”.
4. Add the hydrated salt to the water and seal the calorimeter with its lid. To dissolve the salt, stir the solution but gently.
5. Observe the thermometer, and after being confident that all the mass is dissolved, measure the temperature as it settles off. Record this temperature as “$T_{final,2}$”.
6. Record your experimental measurements on its correspondent table of results.

REPORT OF EXPERIMENT (4)

Title:

Aim:

Results:

	Value
m_{cal} / g	

Reaction 1 (hydrated)		**Reaction 2(anhydrous)**	
	Value		Value
$m_{cal+water,1}$ / g		$m_{cal+water,2}$ / g	
$m_{salt,1}$ / g		$m_{salt,2}$ / g	
$m_{solution,1}$ / g		$m_{solution,2}$ / g	
$T_{initial,1}$ / °C		$T_{initial,2}$ / °C	
$T_{final,1}$ / °C		$T_{final,2}$ / °C	

Calculations:

A) Reaction 1

-Calculate the mass of water:

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

-Calculate the mass of solution:

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

-Calculate the temperature difference:

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

-Calculate the heat exchanged:

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

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

-Calculate the number of moles of the salt:

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

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

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

-Write the thermochemical equation of dissolving the salt:

B) Reaction 2

-Calculate the mass of water:

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

-Calculate the mass of solution:

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

-Calculate the temperature difference:

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

-Calculate the heat exchanged:

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

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

-Calculate the number of moles of the salt:

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

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

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

-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

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

	$CH_3COONa.3H_2O(s)$	$CH_3COONa(s)$	$H_2O(L)$
$\Delta H_{rxn°}$ / kJ mol^{-1}	− 1604.00	− 709.32	− 285.8

$$CH_3COONa(s) + 3H_2O(L) \rightarrow CH_3COONa.3H_2O(s)$$

$$\Delta H_{theor} = \sum \Delta H^{\circ}_{f,products} - \sum \Delta H^{\circ}_{f,reactants}$$

$$\Delta H_{theor} = \{(-1604\ kJ\ )\} - \{(-709.32) + (3 \times -285.8)\} = -37.28\ kJ/mol$$

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:

$$\text{Error\%} = \left|\frac{\Delta H_{theo} - \Delta H_{exp}}{\Delta H_{theo}}\right| \times 100$$

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:

$$3NaOCl + CH_3COCH_3 \rightarrow CH_3COONa + CHCl_3 + 2NaOH$$

- 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), $C_{s,cal}$ is known.
- The experiment involves the determination of the enthalpy of oxidation of acetone by hypochlorite (ΔH_{rxn}).

Substances and materials :

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

50-mL and 250-mL beaker
Thermometer
5-mL and 20-ml graduated pipette
25 mL- graduated cylinder
5.25% (w/v) sodium hypochlorite (Bleach)
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 tube. Record this mass as m_{cal}.
2) Add exactly 10 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 a $T_{initial}$.
4) Add 2.0 mL of 5.0% aqueous acetone into glass tube.
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) Measure the mass of calorimeter tube and the reaction mixture , Record this as $m_{cal+sol}$ 9) Repeat steps (1) – (6) two more times.

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

Calculations

1) Calculation of the mass of solution :

$$m_{sol} = m_{cal+sol} - m_{cal}$$

2) Calculate the total amount of heat that release as a result of reaction "q_{rxn}":

$$q_{rxn} = [(m_{cal} \times C_{s,cal}) + (m_{sol} \times C_{s,sol})] \times (T_{final} - T_{initial})$$

3) 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:

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

4) 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, $n_{acetone}$:

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

5) Knowing the amount of heat (q) evolved per $n_{acetone}$, and according to the pervious reaction equation:

Calculate ΔH_{rxn} in kJ / mole of acetone.: $\Delta H_{rxn} = (- q_{rxn} / n_{acetone})$

REPORT OF EXPERIMENT (5)

Title:

Aim:

Results:

1) Record your measurements in the following table

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

Calculations:

1. Calculation of the mass of solution

2. Calculation the total amount of heat that release as a result of reaction" $q_{rxn} = q_{cal} + q_{sol}$ ":

1. Calculation of the volume of acetone:

2. Calculation the number of acetone moles ($n_{acetone}$)

3. Calculation of ΔH_{rxn} in kJ / mole of acetone:

EXPERIMENT 6: ENTHALPY AND ENTROPY OF FUSION OF WATER

Aim

Understanding that fusion is a change of state of mater 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, ΔH°_{fus}, 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, ΔS°_{fus}, 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:

$$\Delta S^{\circ}_{fus} = \frac{\Delta H^{\circ}_{fus}}{T^{\circ}_{fus}}$$

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 “$m_{1,tube}$”.
  .
- Fill the calorimeter glass tube up to about one third with the hot water from water bath ,temperature that ranges between 40 °C and 50 °C .
- Put the calorimeter glass tube with the hot water inside the Styrofoam. And insert the stirrer and the thermometer.
- Measure the temperature of the the calorimeter glass tube with hot water. Record this temperature as “T_1” Again, weigh the calorimeter glass tube with t hot water. Record this mass as “$m_{2,tube}$”.

- To the hot 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 “T_2”.
- Quickly remove all ice from the calorimeter glass.
- Weigh the calorimeter glass tube with the liquid water inside. Record this mass as “$m_{3,tube}$”.

REPORT OF EXPERIMENT (6)

Title:

Aim:

Results:

Record your measurements in the following table

$m_{1,tube}$ / g	$m_{2,tube}$ / g	$m_{3,tube}$ / g	T_1 / °C	T_2 / °C

$C_{s,water}$ = 4.184 J / g °C
$\Delta H^\circ_{fus,water}$ = 6.01 kJ/mol
$\Delta S^\circ_{fus,water}$ = 22 J/K .mol
$T^\circ_{fus,water}$ = 273K
Mwt_{water} = 18 g/mol

Calculations

$$m_{hot\ water} = m_2 - m_1$$

$$m_{final\ water} = m_3 - m_1$$

$$m_{ice\ melted} = m_3 - m_2$$

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

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

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

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

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

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

-Calculate :

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

-From data given, calculate $\Delta S^{\circ}_{fus,water}$:

$$\Delta S^{\circ}_{fus,water} = \frac{\Delta H^{\circ}_{fus,water}}{T^{\circ}_{fus,water}}$$

-From given data, calculate the percentage of your experimental error in the value of $\Delta S^{\circ}_{fus,water}$ and $\Delta H^{\circ}_{fus,water}$

$$\text{Error\%} = \left| \frac{\Delta S^{\circ}_{actual} - \Delta S^{\circ}_{exp}}{\Delta S^{\circ}_{actual}} \right| \times 100$$

$$\text{Error\%} = \left| \frac{\Delta H^{\circ}_{actual} - \Delta H^{\circ}_{exp}}{\Delta H^{\circ}_{actual}} \right| \times 100$$

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, $Na_2B_4O_5(OH)_4{\cdot}8H_2O$, is a sparingly soluble sodium salt in water. The equation of its solubility in water is presented by the following equation:

$$Na_2B_4O_5(OH)_4{\cdot}8H_2O(s) \xleftrightarrow{H_2O(l)} 2Na^+(aq) + B_4O_5(OH)_4^{-2}(aq) + 8\,H_2O(l)$$

(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:

$$K = \frac{[Na^+(aq)]^2[B_4O_5{}^{-2}(aq)][H_2O(l)]^8}{[Na_2B_4O_5.8H_2O(s)]}$$

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:

$$K \times \frac{[Na_2B_4O_5.8H_2O(s)]}{[H_2O(l)]^8} = [Na_2B_4O_5{}^{-2}(aq)][Na^+(aq)]^2$$

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:

$$\boxed{K_{sp.borax} = [\;B_4O_5{}^{-2}(aq)][Na^+(aq)]^2}$$

- The molarity of borate, $[B_4O_5(OH)_4^{-2}(aq)]$ 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:

$$Na_2B_4O_5(OH)_4^{-2}(aq) + 2HCl(aq) + 3H_2O(l) \rightarrow 4H_3BO_3(aq) + 2Cl^{-}(aq) + 2Na^{+}(aq)$$

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

$$(\frac{M \times V}{n})_{borate} = (\frac{M \times V}{n})_{HCl} \quad \text{So,} \quad M_{borate} = (\frac{M \times V}{n})_{HCl} \times (\frac{n}{V})_{borate}$$

- The chemical equation of borax dissolution given previously shows that the molarity of $Na^{+}(aq)$ is twice the molarity of $Na_2B_4O_5(OH)_4^{2-}(aq)$:

$$[Na^{+}(aq)] = 2 \times [B_4O_5(aq)^{2-}]$$

- Knowing the values of $[B_4O_5(aq)^{2-}]$ and $[Na^{+}(aq)]$ enables us to calculate the value of the solubility product, $K_{sp,borax}$.

- In this experiment the value of $K_{sp,borax}$ will be calculated at two different temperatures (T_1 and T_2). This enables us to evaluate the following thermodynamic parameters:

1) We can calculate the standard change in the Gibbs free energy for dissolving borax in water at the two temperatures using the following equation:

$$\Delta G^{\circ}_{solution} = - R \times T \times \ln K_{sp,borax} \quad , \quad R=8.314 \text{ J/mol. K} \quad , \text{ (T) in kelvin}$$

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

$$\Delta H^{\circ}_{solution} = \frac{\ln \frac{K_{sp,borax,1}}{K_{sp,borax2}} \times R \times T_1 \times T_2}{T_1 - T_2}$$

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

$$\Delta G^{\circ}_{solution} = \Delta H^{\circ}_{solution} - (T \times \Delta S_{solution})$$

This equation can be rearranged to take the following form:

$$\Delta S^{\circ}_{solution} = \frac{\Delta H^{\circ}_{solution} - \Delta G^{\circ}_{solution}}{T}$$

Materials and chemicals

Large for ice and 2 small beakers
Erlenmeyer flasks
Magnetic stirrer or glass rod
Thermometers
A burette filled with standardized HCl solution (about 0.2 mol/L)

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

Procedure

1) In a two small beaker add about 2.5 g of the solid borax and about 50 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 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 "T_1" for the flask at the high temperature and as "T_2" 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 2 drops of methyl orange indicator to each of these two Erlenmeyer flasks. The solution of each should turn from yellow to orange-red.
9) Titrate the sample with the standardize HCl solution. Record the volume of HCl at T_1 as $V_{HCl,1}$ and at T_2 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}.

REPORT OF EXPERIMENT (7)

Title:

Aim:

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 T_1, $[B_4O_5(aq)^{2-}]_1$ and at T_2, $[B_4O_5(aq)^{2-}]_2$:

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

$[B_4O_5(aq)^{2-}]_1$=

$[B_4O_5(aq)^{2-}]_2$=

2. Calculate the molarity of $[Na(aq)^+]$ at T_1, T2

$$[Na(aq)^+]_1 = 2 \times [B_4O_5(aq)^{2-}]_1$$

$$[Na^+(aq)]_2 = 2 \times [B_4O_5(aq)^{2-}]_2$$

3. Using the following equation calculate the value of the solubility product at T_1, $K_{sp,borax,1}$, and at T_2, $K_{sp,borax,2}$,

$$K_{sp,borax,1} = [\text{borate ion}]_1\ [Na+(aq)]_1{}^2$$

$$K_{sp,borax.2} = [\text{borate ion}]_2\ [Na+(aq)]_2{}^2$$

4. Also, using the following equation calculate the value of $\Delta G^{\circ}_{solution}$ at T^1, $\Delta G^{\circ}_{solution,1}$ and at T^2, $\Delta G^{\circ}_{solution,2}$:

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

5. Finally, using the following equation calculate the value of $\Delta H^{\circ}_{solution}$:

$$\Delta H^{\circ}_{solution} = \frac{\ln \frac{K_{sp,borax,1}}{K_{sp,borax2}} \times R \times T_1 \times T_2}{T_1 - T_2}$$

6. Using the following equation calculate the value of $\Delta S^{\circ}_{\text{solution}}$ at T_1, and at T_2:

$$\Delta S^{\circ}_{\text{solution}} = \frac{\Delta H^{\circ}_{\text{solution}} - \Delta G^{\circ}_{\text{solution}}}{T}$$

Tabulate the results of these calculations as follows:

	At "T_1"	At "T_2"
$[B_4O_5(aq)^{2-}]$/mol L^{-1}		
$[Na^+(aq)]$/ /mol L^{-1}		
$K_{sp,borax}$		
$\Delta G_{\text{solution}^{\circ}}$ /kJ mol^{-1}		
$\Delta H^{\circ}_{\text{solution}}$ (kJ/mol)		
$\Delta S^{\circ}_{\text{solution}}$/J K^{-1} mol^{-1}		

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 "$\Delta H^{\circ}_{\text{infin.dilution}}$, and determining the value of $\Delta H_{\text{infin}}{}^{\circ}{}_{\text{.dilution}}$.

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 $\Delta H_{soln}{}^{\circ}$ 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 $\Delta H^{\circ}_{\text{soln}}$ differs as the quantity of the solvent differs. This is due to what is called "enthalpy of dilution.
- This means that the values of $\Delta H^{\circ}_{\text{soln}}$ 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 "$\Delta H^{\circ}_{\text{infin.dilution}}$".
- As mentioned above, the aim of this experiment is to determine "$\Delta H^{\circ}_{\text{infin.dilution}}$.

Materials and chemicals

Balance
Constant-pressure calorimeter
A pipette, a burette or a graduated cylinder
$NH_4Cl(s)$ or $KCl(s)$
Distilled water

Procedure

1) Measure the mass of the glass tube of the constant-pressure calorimeter. Record this mass as "$m_{\text{gass tube}}$".
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 2.5 g of the solute (NH_4Cl). Or (KCl) Record this mass as "m_{solute}".
5) 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 5 g and about 7.5 g of the solid. Record these two masses as "$m_{\text{solute,2}}$" and "$m_{\text{solute,3}}$".

REPORT OF EXPERIMENT (8)

Title:

Aim

Results

Tabulate the results of your experiment and calculations as follows:

	m_{cal} /g	V_{water} /mL	m_{salt}/g	t_1 /°C	t_2 /°C
1					
2					
3					

Density of water, d_{water}, = 1.0 g/mL, Specific heat of solution, $C_{s,soln}$ = specific heat of water, $C_{s,water}$ = 4.184 J/g °C

calorimeter, $C_{s,cal}$=0.836J/g ^{O}C

Mw of NH_4Cl , M_{solute} =53.489 g/mol or Mw (KCl)=74.5 g/mol

Calculations

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

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

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

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

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:

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

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

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

6) For each of the three trials, use the following equation to calculate the heat released per mole of the solute, ΔH°_{soln}:

$$\Delta H^{\circ}_{soln} = \frac{-\,q_{gained}}{n_{solute}}$$

7) Assuming that the volume of each solution, V_{soln} (L), equals the volume of water,

$$V_{soln}\ (L) = V_{water} = 0.05\ L$$

using the following equation, calculate the molarity, M, of each solution:

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

Table of results of calculations:

	m_{soln} /g	Δt /°C	q /J	n_{solute} /mol	ΔH /kJ mol^{-1}	Molarity/ mol L^{-1}
1						
2						
3						

8) Using a graph paper draw the relation between ΔH°_{soln} 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 $\Delta H_{infin^{\circ}.dilution}$:

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

EXPERIMENT (9): Le Chatelier's principle: The common ion effect

Aim

The aim of this experiment t is to understand how changing concentration affects the reversible chemical reactions in terms of their equilibrium constants and positions.

Introduction

- When a sparingly soluble salt is dissolved in water only slight amount dissolves and the rest settles down as an undissolved precipitate.
- The dissolved part is dissociated into its positive and negative ions, while the undissolved part stays as it is. i.e. undissociated.
- A dynamic equilibrium is established between the dissociated ions and the salt that is not dissociated.
- Calcium hydroxide is an example of sparingly soluble salts, and when added to water its dynamic equilibrium is given by the following equation:

$$\mathrm{Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq)}$$

- The equation says that a dynamic equilibrium is established between the solid $\mathrm{Ca(OH)_2}$ and the aqueous $\mathrm{Ca^{2+}}$ and $\mathrm{OH^-}$ ions.
- The equilibrium constant of the above equilibrium is given as:

$$\text{Equilibrium constant} = \frac{[\mathrm{Ca^{2+}}][\mathrm{OH^-}]^2}{[\mathrm{Ca(OH)_2}]}$$

Because the denominator is a solid its concentration by itself is constant. Therefore:

$$\text{Equilibrium constant} \times [\mathrm{Ca(OH)_2}] = [\mathrm{Ca^{2+}}][\mathrm{OH^-}]^2$$

The value of the constant “Equilibrium constant × $[\mathrm{Ca(OH)_2}]$” is given the name “solubility product” symbolized as “$\mathrm{K_{sp}}$”.

- The solubility product of calcium hydroxide is written as $\mathrm{K_{sp,Ca(OH)_2}}$:

$$\mathrm{K_{sp,Ca(OH)_2}} = [\mathrm{Ca^{2+}}][\mathrm{OH^-}]^2$$

- The molarity of the aqueous hydroxide ions, $(\text{molarity})_{\mathrm{OH^-}}$, can be determined by titration with an aqueous solution of hydrogen ions of known molarity.

$$\mathrm{OH^-(aq) + H^+(aq) \rightarrow H_2O(L)}$$

Using the following equation:

$$(\frac{\text{molarity} \times \text{volume}}{\text{n}})_{\text{OH}^-} = (\frac{\text{molarity} \times \text{volume}}{\text{n}})_{\text{H}^+}$$
$$(\frac{\text{molarity} \times \text{volume}}{1})_{\text{OH}^-} = (\frac{\text{molarity} \times \text{volume}}{1})_{\text{H}^+}$$

$$(\text{molarity} \times \text{volume})_{\text{OH}^-} = (\text{molarity} \times \text{volume})_{\text{H}^+}$$

$$(\text{molarity})_{\text{OH}^-} = \frac{(\text{molarity} \times \text{volume})_{\text{H}^+}}{(\text{volume})_{\text{OH}^-}}$$

- From the above, it appears that:

$$[\text{OH}^-]_{\text{total}} = \frac{(\text{molarity} \times \text{volume})_{\text{H}^+}}{(\text{volume})_{\text{OH}^-}}$$

- The molarity of the aqueous calcium ions, $(\text{molarity})_{\text{Ca}^{2+}}$, can be determined after knowing the total molarity of the aqueous hydroxide ions, $(\text{molarity})_{\text{OH}^-}$.
  - ✓ The equation "Ca(OH)_2(s) ⇌ Ca^{2+}(aq) + 2OH^-(aq)" tempts us to say that:

$$[\text{Ca}^{2+}] = \frac{1}{2} \times [\text{OH}^-]_{\text{total}}$$

Here, we have to be very careful, because this is true only when the source of OH^- is only Ca(OH)_2(s), but for cases were the source of OH^- is Ca(OH)_2(s) in addition to other substance such as NaOH(s), then $[\text{Ca}^{2+}]$ does not equal half $[\text{OH}^-]_{\text{total}}$:

$$[\text{Ca}^{2+}] \neq \frac{1}{2} \times [\text{OH}^-]_{\text{total}}$$

  - ✓ In the cases were the source of OH^- is not only Ca(OH)_2(s) but also other substance such as NaOH(s), molarity of OH^- that are given by only Ca(OH)_2(s), $[\text{OH}^-]_{\text{Ca(OH)}_2}$, is the total molarity of hydroxide ions, $[\text{OH}^-]_{\text{total}}$, minus the molarity of hydroxide ions that are given by the other source which, in this experiment, is sodium hydroxide, $[\text{OH}^-]_{\text{NaOH}}$. The molarity of calcium ions, $[\text{Ca}^{2+}]$, is half $[\text{OH}^-]_{\text{Ca(OH)}_2}$.

Materials and chemicals

4 standard measuring flasks
8 conical flasks
Pipette
Burette
5 Funnels and filter papers
Thermometer
Solid calcium hydroxide,
Sodium hydroxide solution (0.1 molar)
Hydrochloric acid (0.1 molar)
Phenolphthalein

Procedure

1) Prepare your filtering flasks using four of the conical flasks, the four funnels and the filter papers. Label your filtering flasks as “A”, “B”, “C” and “D”. keep them aside.
2) Fill the burette with the 0.1 M HCl. Keep it aside.
3) Using the four measuring flasks prepare the following solutions:

Label of solution	$V_{0.1\,M\,NaOH}$	V_{H_2O}	V_{total}
A	0	100	100
B	10	90	100
C	25	75	100
D	50	50	100

Label these measuring flasks as “A”, “B”, “C” and “D”

4) To each of the above solution in the measuring flasks, add (1.5 g) $Ca(OH)_2$, tightly close the flask mouth and carefully shake it to assure saturation with $Ca(OH)_2$.
(*Note: the amount of solid* $Ca(OH)_2$ *must be as little as possible but large enough that little amount of the solid can be clearly seen precipitated. For more information refer to your instructor.*)
5) Filter each solution into its corresponding filtering flask already left aside.
(*Note: Avoid any dilution during filtration.*)
6) Measure the temperature of each solution. Record temperatures as T_A, T_B, T_C and T_D.
(*Note:* T_A, T_B, T_C *and* T_D *are mostly expected to be the same or almost the same.*)
7) Into a conical flask and using pipette, transfer 25 mL of each filtrate. Add 2 drops of the Phenolphthalein to the content. Then titrate with the hydrochloric acid already in the burette. Record the volumes of the acid as $V_{HCl,A}$, $V_{HCl,B}$, $V_{HCl,C}$ and $V_{HCl,D}$.
8) Clean your glassware, put back in location.

Results

1) Tabulate the volume of HCl as follows:

Solution	A	B	C	D
V_{HCl} / mL				

- Calculation of $[Ca^{2+}]$ is as follows:

  First case: If the solution is made up from only calcium hydroxide:

  ✓ The solution, here, contains aqueous calcium hydroxide which presents as ions:

$$Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$$

  ✓ The aqueous solution contains only dissolved calcium hydroxide.

$$[Ca^{2+}] = \frac{1}{2} \times [OH^-]_{total}$$

  Second case: If the solution is made up from both sodium hydroxide and calcium hydroxide:

- ✓ The solution, here, is different from the above. It contains $\mathrm{Ca^{2+}}$, $\mathrm{Na^+}$ and $\mathrm{OH^-}$.
- ✓ The sources of the hydroxide ions here is not only calcium hydroxide but also sodium hydroxide.
- ✓ The source of the calcium ions is only calcium hydroxide. Therefore,

  $$[\mathrm{Ca^{2+}}] \neq \frac{1}{2} \times [\mathrm{OH^-}]_{\mathrm{total}}$$

  Now, the molarity of hydroxide ions that are provided by calcium hydroxide, $[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2}}$, can be calculated by subtracting the molarity of hydroxide ions that are provided by only sodium hydroxide, $[\mathrm{OH^-}]_{\mathrm{NaOH}}$, from the total molarity of hydroxide ions, $[\mathrm{OH^-}]_{\mathrm{total}}$:

  $$[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2}} = [\mathrm{OH^-}]_{\mathrm{total}} - [\mathrm{OH^-}]_{\mathrm{NaOH}}$$

- ✓ Once $[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2}}$ is calculated, $[\mathrm{Ca^{2+}}]$ can also be calculated:

  $$[\mathrm{Ca^{2+}}] = \frac{1}{2} \times [\mathrm{OH^-}]_{\mathrm{Ca(OH)_2}}$$

- After knowing $[\mathrm{Ca^{2+}}]$ and $[\mathrm{OH^-}]_{\mathrm{total}}$ for both of the previous two cases, the solubility product of solid calcium hydroxide, $K_{\mathrm{sp,Ca(OH)_2}}$, can be calculated:

  $$K_{\mathrm{sp,Ca(OH)_2}} = [\mathrm{Ca^{2+}}][\mathrm{OH^-}]^2_{\mathrm{total}}$$

Calculations

1) For the four solutions, and from the results of titrations, calculate the total molarity of hydroxide ions, $[\mathrm{OH^-}]_{\mathrm{total}}$,:

$$[\mathrm{OH^-}]_{\mathrm{total}} = \frac{(\text{molarity} \times \text{volume})_{\mathrm{H^+}}}{(\text{volume})_{\mathrm{OH^-}}}$$

(Note 1: In solution A, $[OH^-]_{total}$ is the molarity of OH^- which is produced only by $Ca(OH)_2$.)
(Note 2: In solution B, C, and D, $[OH^-]_{total}$ is the molarity of OH^- which is produced by both $Ca(OH)_2$ and NaOH.)

2) For the four solutions, and from the information in the table, calculate the molarity of hydroxide ions provided by only NaOH, $[\mathrm{OH^-}]_{\mathrm{NaOH}}$:

$$[\mathrm{OH^-}]_{\mathrm{NaOH}} \times V_{\mathrm{final}} = [\mathrm{OH^-}]_{\mathrm{original}} \times V_{\mathrm{original}}$$

$$[\mathrm{OH^-}]_{\mathrm{NaOH}} \times 100\ \mathrm{mL} = 0.1\ \frac{\mathrm{mol}}{\mathrm{L}} \times V_{\mathrm{original}}$$

(Note: In solution A, no hydroxide ions are provided by NaOH.)

3) For the four solutions, and from the information from the previous calculations, calculate the molarity of hydroxide ions provided by only $Ca(OH)_2$:

$$[OH^-]_{Ca(OH)_2} = [OH^-]_{total} - [OH^-]_{NaOH}$$

4) For the four solutions, calculate the molarity of calcium ions, $[Ca^+]$:

$$[Ca^{2+}] = \frac{1}{2} \times [OH^-]_{Ca(OH)_2}$$

5) Calculate the solubility product of calcium hydroxide, $K_{sp,Ca(OH)_2}$:

$$K_{sp,Ca(OH)_2} = [Ca^{2+}][OH^-]^2_{total}$$

6) Tabulate the results of your calculations as follows:

Solution	A	B	C	D
$[OH^-]_{total}$ / mol L^{-1}				
$[Ca^+]$ / mol L^{-1}				
$K_{sp,Ca(OH)_2}$				

Questions

1) How does the addition of more hydroxide ions to the saturated solution of $Ca(OH)_2$ affect the position of equilibrium between the dissolved $Ca(OH)_2$ and the undissolved $Ca(OH)_2$?
2) How does the addition of more hydroxide ions to the saturated solution of $Ca(OH)_2$ affect the value of its K_{sp}?
3) Knowing that the literature value of $K_{sp,Ca(OH)_2}$ equals 5.5×10^{-6}, calculate your experimental error.

REPORT OF EXPERIMENT (9)

Title: Le Chatelier's principle: The common ion effect

Aim: The aim of this experiment is to understand how changing concentration affects the reversible chemical reaction in terms of its equilibrium constant and position.

Results:

Solution	A	B	C	D
V_{HCl} / mL				

Calculations

1) The total molarity of hydroxide ions, $[\mathrm{OH}^-]_{\mathrm{total}}$:

A) $[\mathrm{OH}^-]_{\mathrm{total,A}} = \frac{(\mathrm{molarity} \times \mathrm{volume})_{\mathrm{H}^+}}{(\mathrm{volume})_{\mathrm{OH}^-}}$

B) $[\mathrm{OH}^-]_{\mathrm{total,B}} = \frac{(\mathrm{molarity} \times \mathrm{volume})_{\mathrm{H}^+}}{(\mathrm{volume})_{\mathrm{OH}^-}}$

C) $[\mathrm{OH}^-]_{\mathrm{total,C}} = \frac{(\mathrm{molarity} \times \mathrm{volume})_{\mathrm{H}^+}}{(\mathrm{volume})_{\mathrm{OH}^-}}$

D) $[\mathrm{OH}^-]_{\mathrm{total,D}} = \frac{(\mathrm{molarity} \times \mathrm{volume})_{\mathrm{H}^+}}{(\mathrm{volume})_{\mathrm{OH}^-}}$

$[\mathrm{OH}^-]_{\mathrm{total,D}}$ = The molarity of hydroxide ions provided by NaOH, $[\mathrm{OH}^-]_{\mathrm{NaOH}}$:

A) $[\mathrm{OH}^-]_{\mathrm{NaOH,A}} \times V_{\mathrm{A}} = [\mathrm{OH}^-]_{\mathrm{original}} \times V_{\mathrm{original}}$

$[\mathrm{OH}^-]_{\mathrm{NaOH,A}} = \frac{[\mathrm{OH}^-]_{\mathrm{original}} \times V_{\mathrm{original,A}}}{V_{\mathrm{A}}}$

B) $(\mathrm{Molarity}_{\mathrm{OH}^-})_{\mathrm{B}} \times V_{\mathrm{B}} = (\mathrm{Molarity}_{\mathrm{OH}^-})_{\mathrm{original}} \times V_{\mathrm{original}}$

$[\mathrm{OH}^-]_{\mathrm{NaOH,B}} = \frac{[\mathrm{OH}^-]_{\mathrm{original}} \times V_{\mathrm{original,A}}}{V_{\mathrm{B}}}$

C) $[\mathrm{OH^-}]_{\mathrm{NaOH,C}} \times \mathrm{V_C} = (\mathrm{Molarity_{OH^-}})_{\mathrm{original}} \times \mathrm{V_{original}}$

$$[\mathrm{OH^-}]_{\mathrm{NaOH,C}} = \frac{(\mathrm{Molarity_{OH^-}})_{\mathrm{original}} \times \mathrm{V_{original,C}}}{\mathrm{V_C}}$$

D) $[\mathrm{OH^-}]_{\mathrm{NaOH,D}} \times \mathrm{V_D} = (\mathrm{Molarity_{OH^-}})_{\mathrm{original}} \times \mathrm{V_{original}}$

$$[\mathrm{OH^-}]_{\mathrm{NaOH,D}} = \frac{(\mathrm{Molarity_{OH^-}})_{\mathrm{original}} \times \mathrm{V_{original,D}}}{\mathrm{V_D}}$$

2) The molarity of hydroxide ions provided by only $\mathrm{Ca(OH)_2}$, $[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2}}$:

A) $[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,A}} = [\mathrm{OH^-}]_{\mathrm{total,A}} - [\mathrm{OH^-}]_{\mathrm{NaOH,A}}$

B) $[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,B}} = [\mathrm{OH^-}]_{\mathrm{total,B}} - [\mathrm{OH^-}]_{\mathrm{NaOH,B}}$

C) $[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,C}} = [\mathrm{OH^-}]_{\mathrm{total,C}} - [\mathrm{OH^-}]_{\mathrm{NaOH,C}}$

D) $[\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,D}} = [\mathrm{OH^-}]_{\mathrm{total,D}} - [\mathrm{OH^-}]_{\mathrm{NaOH,D}}$

3) The molarity of calcium ions, $[\mathrm{Ca^{2+}}]$:

A) $[\mathrm{Ca^{2+}}]_{\mathrm{A}} = = \frac{1}{2} \times [\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,A}}$

B) $[\mathrm{Ca^{2+}}]_{\mathrm{B}} = = \frac{1}{2} \times [\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,B}}$

C) $[\mathrm{Ca^{2+}}]_{\mathrm{C}} = = \frac{1}{2} \times [\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,A}}$

D) $[\mathrm{Ca^{2+}}]_{\mathrm{D}} = = \frac{1}{2} \times [\mathrm{OH^-}]_{\mathrm{Ca(OH)_2,A}}$

4) The solubility product, $\mathrm{K_{sp,Ca(OH)_2}}$. $\mathrm{K_{sp,Ca(OH)_2}} = [\mathrm{Ca^{2+}}][\mathrm{OH^-}]^2_{\mathrm{total}}$

A) $\mathrm{K_{sp,Ca(OH)_2,A}} = [\mathrm{Ca^{2+}}]_{\mathrm{A}}[\mathrm{OH^-}]^2_{\mathrm{total,A}}$

B) $K_{sp,Ca(OH)_2,B} = [Ca^{2+}]_B[OH^-]^2_{total,B}$

C) $K_{sp,Ca(OH)_2,C} = [Ca^{2+}]_C[OH^-]^2_{total,C}$

D) $K_{sp,Ca(OH)_2,D} = [Ca^{2+}]_D[OH^-]^2_{total,D}$

5) Table of the results of calculations:

Solution	A	B	C	D
$[OH^-]$ / mol L^{-1}				
$[Ca^{2+}]$ / mol L^{-1}				
$K_{sp,Ca(OH)_2} \times 10^{-5}$				

Answers of the questions

1)

2)

3) The experimental error:

Experimental error = $\pm \frac{(K_{sp,Ca(OH)_2})_{literature} - (K_{sp,Ca(OH)_2})_{experimental}}{(K_{sp,Ca(OH)_2})_{literature}} \times 100$

A)

B)

C)

D)

EXPERIMENT (10): SOLUBILITY PRODUCT AND THERMODYNAMIC FUNCTIONS

Aim

The aim is obtaining the change in Gibbs free energy for the solubility of salts.

Introduction

- The relationship between standard change in Gibb free energy, $\Delta G°$, change in enthalpy, $\Delta H°$, and standard change in entropy, $\Delta S°$, is given as:

$$\Delta G° = \Delta H° - T\Delta S°$$

- The relationship between the standard change in Gibb free energy, $\Delta G°$, and the equilibrium constant is given as:

$$\Delta G° = - R\,T\ln K$$

- From the above two equations, the following equation is obtained:

$$\ln K = \frac{-\Delta H°}{R} \times \frac{1}{T} + \frac{\Delta S°}{R}$$

- The "solubility product", K_{sp},.is an equilibrium constant which expresses the equilibrium established between a solid salt and its dissociated ions in a saturated solution of the salt.
- As any other equilibrium constant, K_{sp} is affected by temperature.
- Examples are:
  - ✓ Sodium chloride, NaCl:

$$NaCl(s) \xrightarrow{water} Na^+(aq) + Cl^-(aq)$$

$$K_{sp,NaCl}, = [Na^+]\,[Cl^-]$$

  - ✓ Copper sulphate, $CuSO_4$:

$$CuSO_4 \cdot 5H_2O(s) \xrightarrow{water} Cu^{2+}(aq) + SO_4^{2-}(aq) + 5H_2O(L)$$

$$K_{sp,CuSO_4} = [Cu^{2+}]\,[SO_4^{2-}]$$

- As “y = m x + b” is a straight-line equation the equation “$\ln K_{sp} = \frac{-\Delta H°}{R} \times \frac{1}{T} + \frac{\Delta S°}{R}$” also is a straight-line equation. Plot of “$\ln K_{sp}$” versus “$\frac{1}{T}$” gives a straight line in which:

$$\text{slope} = \frac{-\Delta H°}{R}$$

$$\text{intercept} = \frac{\Delta S°}{R}$$

- ΔG°, which usually cannot be determined experimentally, but can be obtained after experimental determination of both ΔH° and ΔS°.
- From literature, obtain the value of both ΔH° and ΔS° for each salt and use these values to calculate the experimental error:

$$\text{experimental error} = \pm \frac{\Delta H^{\circ}_{\text{literature}} - \Delta H^{\circ}_{\text{experimental}}}{\Delta H^{\circ}_{\text{literature}}} \times 100$$

$$\text{experimental error} = \pm \frac{\Delta S^{\circ}_{\text{literature}} - \Delta S^{\circ}_{\text{experimental}}}{\Delta S^{\circ}_{\text{literature}}} \times 100$$

Materials and chemicals

6 beakers (250 mL)
6 beakers (50 mL)
3 Pipettes (5 mL)
3 Thermometer
Solid sodium chloride
Solid copper sulfate pentahydrate
Balance
Two thermostats (one is peradjusted at 40 °C and the second one at 60 °C)
Hotplates or oven at 150 °C

Procedure

Before starting, prepare two water baths one at approximately 40 °C and 60 °C.

salt,1A is NaCl at room temperature salt,1B is NaCl at 40 °C salt,1C is NaCl at 60 °C
salt,2A is $CuSO_4{\cdot}5H_2O$ at room temperature salt,2B is $CuSO_4{\cdot}5H_2O$ at 40 °C salt,2C is $CuSO_4{\cdot}5H_2O$ at 60 °C

1) Label three 50-mL beakers as “salt,1A”, “salt,1B” and “salt,1C”. Cover it with a proper lid then weigh each and record masses as $m_{\text{initial,1A}}$, , $m_{\text{initial,1B}}$ and $m_{\text{initial,1C}}$. Leave these beakers aside.

2) Label three 50-mL beakers as “salt,2A”, “salt,2B” and “salt,2C”. Cover it with a proper lid then weigh each and record masses as $m_{\text{initial,2A}}$, , $m_{\text{initial,2B}}$ and $m_{\text{initial,2C}}$. Leave these beakers aside.
3) Label three 250-mL beakers as “salt,1A”, “salt,1B” and “salt,1C”; then add more than 5 g of NaCl(s) and 10 mL of distilled water in each beaker **to assure saturation**.
4) Label three 250-mL beakers as “salt,2A”, “salt,2B” and “salt,2C”; then add 5g of salt and 10 mL of distilled water in each beaker **to assure saturation**.
5) Leave the 250-mL beakers labeled salt,1A and salt,2A aside at room temperature.
6) Put the 250-mL beakers labeled salt,1B and salt,2B inside the 40 °C-water bath.
7) Put the 250-mL beakers labeled salt,1C and salt,2C inside the 60 °C-water bath.
8) For at least 15 minutes, stir continuously each of the six beakers by a glass stirrer.
9) Use a proper pipette to transfer 5 mL from each solution in each of the six 250 mL beakers to each of its correspondent six 50-mL beaker.
   (*Note: After each transfer, be sure that your pipette is rinsed with water at the same temperature of the water bath three times before used for the next transfer*)
10) Put each of the six 50-mL beakers in an oven or on hot plates for as long as it takes to ensure total vaporization of water and total dryness.
    (*Note: Be so careful while heating that no splashes occurs otherwise experiment will be seriously affected.*)
11) Weigh each of the six 50-mL beakers. Record masses as $m_{\text{final,1A}}$, , $m_{\text{final,1B}}$ and $m_{\text{final,1C}}$ and $m_{\text{final,2A}}$, , $m_{\text{final,2B}}$ and $m_{\text{final,2C}}$
12) Clean your all glassware, put back in location.

Results

Tabulate your experimental results as follows:

$m_{\text{initial,1A}}$ /g	$m_{\text{initial,1B}}$ /g	$m_{\text{initial,1C}}$ /g

$m_{\text{initial,2A}}$ /g	$m_{\text{initial,2B}}$ /g	$m_{\text{initial,2C}}$ /g

$m_{\text{final,1A}}$ /g	$m_{\text{final,1B}}$ /g	$m_{\text{final,1C}}$ /g

$m_{\text{final,2A}}$ /g	$m_{\text{final,2B}}$ /g	$m_{\text{final,2C}}$ /g

Calculations

1) After complete evaporation at each temperature, calculate the mass of the residue, m_{residue},:
   - ✓ Mass of the residues of salt number 1 at temperatures A, B and C:

$$m_{\text{1,residues,A}} = m_{\text{final,1A}} - m_{\text{initial,1A}}$$

$$m_{\text{1,residues,B}} = m_{\text{final,1B}} - m_{\text{initial,1B}}$$

$m_{1,\text{residues},C} = m_{\text{final},1C} - m_{\text{initial},1C}$

- ✓ Mass of the residues of salt number 2 at temperatures A, B and C:

$m_{2,\text{residues},A} = m_{\text{final},2A} - m_{\text{initial},2A}$

$m_{2,\text{residues},B} = m_{\text{final},2B} - m_{\text{initial},2B}$

$m_{2,\text{residues},C} = m_{\text{final},2C} - m_{\text{initial},2C}$

2) Calculate the solubility in mol/L units:

$$s = \frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{\text{g}}{\text{mol}})}}{\text{Volume of solution (L)}}$$

(*Note: the volume used in this equation is the volume of the saturated solutions taken to be dried*)

- ✓ The solubility of salt number 1 at temperatures A, B and C are abbreviated as follows:

$S_{1,A} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{\text{g}}{\text{mol}})}}{\text{Volume of solution (L)}}\right)_{1,A} =$

$S_{1,B} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{\text{g}}{\text{mol}})}}{\text{Volume of solution (L)}}\right)_{1,B} =$

$S_{1,C} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{\text{g}}{\text{mol}})}}{\text{Volume of solution (L)}}\right)_{1,C} =$

- ✓ The solubilities of salt number 2 at temperatures A, B and C are abbreviated as follows:

$S_{2,A} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{\text{g}}{\text{mol}})}}{\text{Volume of solution (L)}}\right)_{2,A} =$

$S_{2,B} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{\text{g}}{\text{mol}})}}{\text{Volume of solution (L)}}\right)_{2,B} =$

$S_{2,C} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{\text{g}}{\text{mol}})}}{\text{Volume of solution (L)}}\right)_{2,C} =$

3) Tabulate your results of the previous calculations as follows:

Temperature		A	B	C
Salt 1	m_{residue}			
	s			

Salt 2	$m_{residue}$			
	s			

4) Calculate the solubility products:

✓ The solubility product of salt number 1 at temperatures A, B and C:

$K_{sp,1,A} = S_{1,A}^2 =$

$K_{sp,1,B} = S_{1,B}^2 =$

$K_{sp,1,C} = S_{1,C}^2 =$

✓ The solubility product of salt number 2 at temperatures A, B and C:

$K_{sp,2,A} = S_{2,A}^2 =$

$K_{sp,2,B} = S_{2,B}^2 =$

$K_{sp,2,C} = S_{2,C}^2 =$

5) Calculate the values of "$\frac{1}{T}$" and "ln $K_{sp,salt}$" and tabulate as follows:

	A	B	C
$\frac{1}{T}$ (K^{-1})			
$\ln K_{sp,salt,1}$			
$\ln K_{sp,salt,2}$			

6) Draw the relation given in the equation "$\ln K_{sp} = \frac{-\Delta H°}{R} \times \frac{1}{T} + \frac{\Delta S°}{R}$".

7) From the graph:

✓ **Salt number 1**

Obtain slope_1, then calculate ΔH_1°:

$\text{slope}_1 = \frac{-\Delta H_1^\circ}{R}$

$\Delta H_1^\circ = -\,\text{slope}_1 \times R$

Obtain intercept_1, then calculate ΔS_1°:

$\text{intercept}_1 = \frac{\Delta S_1^\circ}{R}$

$\Delta S_1^\circ = \text{intercept}_1 \times R$

✓ **Salt number 2**

Obtain slope_2, then calculate ΔH_2°:

$\text{slope}_2 = \frac{-\Delta H_2^\circ}{R}$

$\Delta H_2^{\circ} = -\text{slope}_2 \times R$

Obtain intercept_2, then calculate ΔS_2°:

$\text{intercept}_2 = \frac{\Delta S_2^{\circ}}{R}$

$\Delta S_2^{\circ} = \text{intercept}_2 \times R$

8) Calculate $\Delta G°$ and tabulate as follows
$\Delta G° = \Delta H° - T \times \Delta S°$

	25 °C	40 °C	60 °C
ΔG_1°			
ΔG_2°			

9) Calculate the experimental error in both $\Delta H°$ and $\Delta S°$:

$$\text{experimental error} = \pm \frac{\Delta H^{\circ}_{\text{literature}} - \Delta H^{\circ}_{\text{experimental}}}{\Delta H^{\circ}_{\text{literature}}} \times 100$$

$$\text{experimental error} = \pm \frac{\Delta S^{\circ}_{\text{literature}} - \Delta S^{\circ}_{\text{experimental}}}{\Delta S^{\circ}_{\text{literature}}} \times 100$$

REPORT OF EXPERIMENT (10)

Title: solubility product and thermodynamic functions

Aim: Obtaining the change in Gibbs free energy for the solubility of salts.

Results:

$m_{initial,1A}$ /g	$m_{initial,1B}$ /g	$m_{initial,1C}$ /g

$m_{initial,2A}$ /g	$m_{initial,2B}$ /g	$m_{initial,2C}$ /g

$m_{final,1A}$ /g	$m_{final,1B}$ /g	$m_{final,1C}$ /g

$m_{final,2A}$ /g	$m_{final,2B}$ /g	$m_{final,2C}$ /g

Calculations

1) Mass of the residues:

- ✓ salt number 1 at temperatures A, B and C:

$m_{1,residues,A} = m_{final,1A} - m_{initial,1A} =$

$m_{1,residues,B} = m_{final,1B} - m_{initial,1B} =$

$m_{1,residues,C} = m_{final,1C} - m_{initial,1C} =$

- ✓ salt number 2 at temperatures A, B and C:

$m_{2,residues,A} = m_{final,2A} - m_{initial,2A} =$

$m_{2,residues,B} = m_{final,2B} - m_{initial,2B} =$

$m_{2,\text{residues},C} = m_{\text{final},2C} - m_{\text{initial},2C} =$

2) The solubility in mol/L units:
   - ✓ The solubility of salt number 1 at temperatures A, B and C are abbreviated as follows:

$$S_{1,A} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{g}{mol})}}{\text{Volume of solution (L)}}\right)_{1,A} =$$

$$S_{1,B} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{g}{mol})}}{\text{Volume of solution (L)}}\right)_{1,B} =$$

$$S_{1,C} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{g}{mol})}}{\text{Volume of solution (L)}}\right)_{1,C} =$$

   - ✓ The solubility of salt number 2 at temperatures A, B and C are abbreviated as follows:

$$S_{2,A} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{g}{mol})}}{\text{Volume of solution (L)}}\right)_{2,A} =$$

$$S_{2,B} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{g}{mol})}}{\text{Volume of solution (L)}}\right)_{2,B} =$$

$$S_{2,C} = \left(\frac{\frac{\text{mass of salt (g)}}{\text{molar mass of salt }(\frac{g}{mol})}}{\text{Volume of solution (L)}}\right)_{2,C} =$$

3) Table of results calculations:

Temperature		A	B	C
Salt 1	m_{residue}			
	s			
Salt 2	m_{residue}			
	s			

4) Calculation of the solubility products
   - ✓ The solubility product of salt number 1 at temperatures A, B and C:

$K_{sp,1,A} = S_{1,A}^2 =$

$K_{sp,1,B} = S_{1,B}^2 =$

$K_{sp,1,C} = S_{1,C}^2 =$

   - ✓ The solubility product of salt number 2 at temperatures A, B and C:

$K_{sp,2,A} = S_{2,A}^2 =$

$K_{sp,2,B} = S_{2,B}^2 =$

$K_{sp,2,C} = S_{2,C}^2 =$

5) Calculation of "$\frac{1}{T}$" and "ln $K_{sp,salt}$":

	A	B	C
$\frac{1}{T}$ (K^{-1})			
ln $K_{sp,salt,1}$			
ln $K_{sp,salt,2}$			

6) From the graph:

- ✓ **Salt number 1**

  Obtaining $slope_1$ and calculating ΔH_1°:

  $slope_1 = \frac{-\Delta H_1^\circ}{R}$

  $\Delta H_1^\circ = -\, slope_1 \times R$

  Obtaining $intercept_1$ and calculating ΔS_1°:

  $intercept_1 = \frac{\Delta S_1^\circ}{R}$

  $\Delta S_1^\circ = intercept_1 \times R$

- ✓ **Salt number 2**

  Obtaining $slope_2$ and calculating ΔH_2°:

  $slope_2 = \frac{-\Delta H_2^\circ}{R}$

$\Delta H_2^{\circ} = -\,\text{slope}_2 \times R$

Obtaining intercept_2 and calculating ΔS_2°:

$\text{intercept}_2 = \frac{\Delta S_2^{\circ}}{R}$

$\Delta S_2^{\circ} = \text{intercept}_2 \times R$

7) Calculation and tabulation of ΔG°:

- ✓ **Salt number 1**

$\Delta G_1^{\circ} = \Delta H_1^{\circ} - T\,\Delta S_1^{\circ}$

- ✓ **Salt number 2**

$\Delta G_2^{\circ} = \Delta H_2^{\circ} - T\,\Delta S_2^{\circ}$

ΔG_1°	
ΔG_2°	

8) Calculate the experimental error in both ΔH° and ΔS°:

- ✓ **Salt number 1**

$$\text{experimental error}_{\Delta H_1^{\circ}} = \pm \frac{\Delta H^{\circ}_{\text{literature},1} - \Delta H^{\circ}_{\text{experimental},1}}{\Delta H^{\circ}_{\text{literature},1}} \times 100$$

$$\text{experimental error}_{\Delta S_1^\circ} = \pm \frac{\Delta S^\circ_{\text{literature},1} - \Delta S^\circ_{\text{experimental},1}}{\Delta S^\circ_{\text{literature},1}} \times 100$$

✓ **Salt number 2**

9) $$\text{experimental error}_{\Delta H_2^\circ} = \pm \frac{\Delta H^\circ_{\text{literature},2} - \Delta H^\circ_{\text{experimental},2}}{\Delta H^\circ_{\text{literature},2}} \times 100$$

10) $$\text{experimental error}_{\Delta S_2^\circ} = \pm \frac{\Delta S^\circ_{\text{literature},2} - \Delta S^\circ_{\text{experimental},2}}{\Delta S^\circ_{\text{literature},2}} \times 100$$

CHEM 232
PRELAB QUESTIONS

Experiment (1):

1) In the procedure, what are the differences between systems A, system B and system C.

2) How many times you will use the thermometer.

3) What is the condition under which you should comply with before recording temperature?

Pre-lab questions

Experiment (2)

1) Define calorimetry.

2) What are the constituent of the constant-pressure calorimeter?

3) What is the substance which loses heat and that which gains this heat?

4) If 0.836 J/g °C is the literature value of the specific heat of the calorimeter tube and $C_{s,cal}$ is the one you determined in your experiment, so how would you calculate your experimental error percentage . (Write the law).

Experiment (3)

1) Write the equation used to calculate the amount of heat exchanged as a result of changing the temperature of a substance.

2) Why should you stir the content of the calorimeter tube before you measure and record its temperature?

3) Why transferring the heated metal from the hot water to the calorimeter should be done quickl

Pre-lab questions

Experiment (4):

1) In this experiment what is the enthalpy of hydration?

2) In step number 5 of the procedure you are asked to be confident about something. What is that thing? And why should you be confident about it.

3) After calculating(q) , explain how will you calculate (ΔH) ?

Experiment (5) :

1) Write the chemical equation of the oxidation of acetone by sodium hypochlorite.

2) Why is it wrong to stir by thermometer?

3) How can you calculate the volume of acetone?

4) In calculations, why the sodium hypochlorite is not used instead of the acetone?

Pre-lab questions

Experiment (6):

1) Define the standard enthalpy of fusion .

2) In the procedure, step number 10 says:
   "Make sure that the mixture contains water in both its liquid state and its solid state" Why?

3) If you are asked to determine $\Delta H^{\circ}_{\text{fus,water}}$, why its value is given in the data of the experiment?

Experiment (7):

1) The equation of the solubility of borax in water is:

   $Na_2B_4O_5(OH)_4{\cdot}8H_2O(s) \xleftrightarrow{H_2O(l)} 2Na^+(aq) + B_4O_5(OH)_4^{2-}(aq) + 8\ H_2O(l)$

   While the equilibrium constant is given by the following equation:

   $K_{sp,borax} = [Na_2B_4O_5(aq)^{2-}][Na^+(aq)]^2$

   Why did not $Na_2B_4O_5(OH)_4{\cdot}8H_2O(s)$ and $H_2O(l)$ appear in the equilibrium constant equation?

2) Calculations requires the concentration of both borate and sodium ions. However, in the procedure you only are able to determine the concentration of borate ions. How can you evaluate the concentration of the sodium ions?

3) What is the indicator used for titration? How does it color change?

Pre-lab questions

Experiment (8):

1) What is enthalpy of solution at infinite dilution?

2) In the procedure, which step is the one that may cause more experimental failure? Rationalize.

3) Write the equation used to calculate (q).

Experiment (9):

1) Why is added mass of solid $Ca(OH)_2$ is critical ?

2) What would the color of the solution in the conical flask be after adding some drops of (ph.ph) indicator ? And what would the color be at the end of titration?

3) In the step number (5) of the procedure you were asked to avoid dilution with water? What is the result if you do not comply?

Pre-lab questions

Experiment (10):

1) In procedure steps number 3 and 4 you are asked to dissolve about (5 g)of salt in 10 ml of distilled water . Why the added mass of solid is critical ?

2) In procedure step (10), Why would splashing out the solution affect the accuracy of your experimint ?

3) Write the equation you will use to draw the required graph. Mention the slope and the intercept of this equation .