Name_____Lab Day_____Lab Time_____

Experiment 6 · Calorimetry

Pre-lab questions
Answer these questions and hand them to the TF before beginning work.

(1) What is the purpose of this experiment?

_____

_____

(2) How is heat q related to enthalpy change ΔH at constant pressure?

_____

_____

(3) You will determine the calorimeter constant $\mathscr{C}_{cal}$ of your calorimeter. What property of a calorimeter does $\mathscr{C}_{cal}$ measure?

_____

_____

(4) Suppose you determine an enthalpy change at constant pressure to be negative. Is heat absorbed or released?

_____

_____

(5) You will calculate several temperature changes ΔT using the formula $\Delta T = T_{mix} - T_i$. How will you determine the value of T_{mix}?

_____

_____

Calorimetry

Mathematical development

The calorimeter constant $\mathscr{C}_{cal}$

Calorimetry is the science of measuring the quantities of heat released or absorbed during a chemical reaction. The amount of heat that flows in or out of a system depends on (1) the quantity of matter that constitutes the system, (2) the identity of that matter, and (3) temperature change experienced by the system as it absorbs or releases heat. In equation form

$$q = mC\Delta T$$

where q represents heat, m is the mass of the system, C is the heat capacity of the system, and ΔT is the temperature change. The heat capacity C is a measure of how a substance responds to the absorption or release of heat; substances that have a low value of C such as iron (C = 0.45 J/(g·°C)) tend to be good conductors of heat whereas substances that have a high value of C such as water (C = 4.18 J/(g·°C)) tend to be good insulators.

A vessel called a calorimeter (Figure 6–1) is needed to hold the substances under study. Ideally, the calorimeter is a perfect insulator, that is, it should prevent the loss of heat from the system to the surroundings and it should not allow heat from the surroundings to enter the system. In practice, this state of affairs is extremely difficult to achieve: heat is inevitably exchanged between the system and the surroundings.

Consider what happens when a quantity of hot water is poured into a quantity of cold water inside a calorimeter. The following relationship accounts for the heat exchanged:

heat lost by   =   heat gained by   +   heat gained by
the hot water   the cold water   the calorimeter

This relationship is expressed symbolically as

$$-q_{hw} = q_{cw} + q_{cal}$$
$$-m_{hw}C_{hw}\Delta T_{hw} = m_{cw}C_{cw}\Delta T_{cw} + \mathscr{C}_{cal}\Delta T_{cw} \qquad \text{(Eqn. 6-1)}$$

where m_{hw} is the mass of the hot water, m_{cw} is the mass of the cold water, C_{hw} is the heat capacity of the hot water, C_{cw} is the heat capacity of the cold water, ΔT_{hw} is the temperature change experienced by the hot water as a result of mixing with the cold water, ΔT_{cw} is the temperature change experienced by the cold water as a result of mixing with the hot water, and $\mathscr{C}_{cal}$ denotes the calorimeter constant, which is a measure of the calorimeter's ability to act as an insulator.

We now employ the substitutions

$$m_{hw} = \rho_{hw}V_{hw} \qquad \text{and} \qquad m_{cw} = \rho_{cw}V_{cw}$$

where ρ represents density and V represents volume. For the sake of simplicity we ignore the fact that the density and the heat capacity of water vary with temperature; thus,

$$\rho_{cw} = \rho_{hw} = \rho_w \qquad \text{and} \qquad C_{cw} = C_{hw} = C_w$$

Plugging these substitutions into Eqn. 6-1 and solving for $\mathscr{C}_{cal}$ gives

$$\mathscr{C}_{cal} = -\rho_w C_w \left[V_{hw}\left(\frac{\Delta T_{hw}}{\Delta T_{cw}}\right) + V_{cw} \right]$$

Inserting the numerical values ρ_w = 1.00 g/mL and C_w = 4.18 J/(g·°C) gives
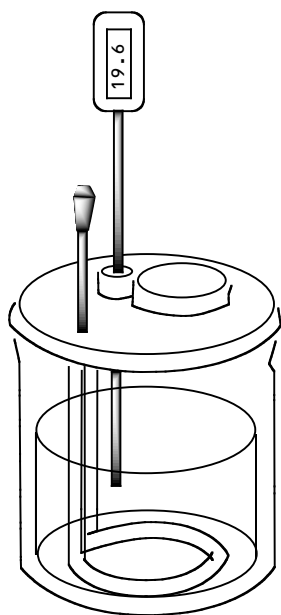

Figure 6-1 A see-through view of the calorimeter. An insulated inner cup holds the solutions being studied. The lid of the vessel is equipped with a cap (which should remain in place at all times), a thermometer, and a stir-ring ring.

$$\mathscr{C}_{cal} = -\left(\frac{4.18\text{ J}}{\text{mL}\cdot{}^{\circ}\text{C}}\right)\left[V_{hw}\left(\frac{\Delta T_{hw}}{\Delta T_{cw}}\right) + V_{cw}\right] \qquad \text{(Eqn. 6-2a)}$$

You use a slightly modified version of Eqn. 6-2a to compute $\mathscr{C}_{cal}$ of your calorimeter.

The molar enthalpy change of reaction $\Delta\mathscr{H}$

At constant pressure, the heat in or out of a system is equal to the enthalpy change ΔH. No special arrangement is needed to guarantee that the experiment takes place at constant pressure: the atmosphere automatically provides a constant pressure of 1 atm.

Suppose that the system consists of a salt MB(s) that absorbs heat when it dissolves in water in a calorimeter:

$$MB(s) + \text{heat} \rightarrow MB(aq)$$

The following relationship accounts for the heat exchanged:

heat gained by	=	heat lost by	+	heat lost by
the system		the solution		the calorimeter
(i.e., by the salt)				

At constant pressure, this heat equation can be expressed as

$$q_{sys} = -q_{soln} - q_{cal}$$
$$\Delta H_{sys} = -m_{soln}C_{soln}\Delta T - \mathscr{C}_{cal}\Delta T \qquad \text{(Eqn. 6-3a)}$$

where ΔH_{sys} is the enthalpy change of the system (i.e., the salt), m_{soln} is the mass of the solution, C_{soln} is the heat capacity of the solution, ΔT is the temperature change experienced by the solution as a result of the dissolution of the salt, and $\mathscr{C}_{cal}$ is the calorimeter constant.

To make the measurement of ΔH_{sys} more meaningful, it is customary to report its value on a per-mole basis. Thus, we are really interested in $\Delta H_{sys}/n_{sys}$, where n_{sys} is the number of moles of salt that dissolve. We will give the quantity $\Delta H_{sys}/n_{sys}$ the symbol $\Delta\mathscr{H}$. Because the number of moles n of a substance is related to the mass m of that substance and to its molar mass $\mathscr{M}$ by

$$n = m/\mathcal{M}$$

dividing both sides of Eqn. 6-3a by n_{sys} and making the substitutions $\Delta H_{sys}/n_{sys} = \Delta\mathcal{H}$ and $n_{sys} = m_{sys}/\mathcal{M}_{sys}$ gives

$$\Delta\mathcal{H} = -\frac{m_{soln}C_{soln}\Delta T + \mathscr{C}_{cal}\Delta T}{m_{sys}/\mathcal{M}_{sys}} \qquad \text{(Eqn. 6-4)}$$

As before, the mass m_{soln} of the salt solution can be related to its density and its volume:

$$m_{soln} = \rho_{soln}V_{soln}$$

For simplicity, we assume that the aqueous salt solution is so dilute that its density ρ_{soln} and heat capacity C_{soln} are the same as those of pure water, that is, $\rho_{soln} = \rho_w = 1.00$ g/mL and $C_{soln} = C_w = 4.18$ J/(g·°C). Substituting these values into Eqn. 6-4 gives Eqn. 6-5:

$$\Delta\mathcal{H} = -\left(\frac{\mathcal{M}_{sys}\Delta T}{m_{sys}}\right)\left[\left(\frac{4.18\text{ J}}{\text{mL·°C}}\right)V_{soln} + \mathscr{C}_{cal}\right] \qquad \text{(Eqn. 6-5)}$$

You will use Eq. 6-5 to compute the molar enthalpy change $\Delta\mathcal{H}$ of dissolving various salts in water.

Thermodynamics of solution

In principle the process of dissolving a salt MB(s) can be deconstructed into three steps (Figure 6-2). First, energy from the surroundings must be invested to completely overcome the strong ion–ion intermolecular forces that are attracting the positive M$^+$ ions to the negative B$^-$ ions:

$$\text{MB(s)} \rightarrow \text{M}^+\text{(g)} + \text{B}^-\text{(g)} \qquad \Delta H_{latt,MB} > 0$$
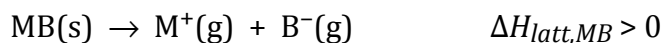
This quantity of energy is called the lattice energy $\Delta H_{latt,MB}$ of the salt MB: it is a positive number because energy must be added to the salt in order to completely separate its constituent ions.

Next, the $M^+(g)$ ions are surrounded by water molecules and enter solution:

$$M^+(g) \rightarrow M^+(aq) \qquad \Delta H_{soln,M+} < 0$$

This quantity of energy is called the enthalpy of solution of the M^+ ion $\Delta H_{soln,M+}$: it is a negative number because the energy of the system drops as the positive charge of M^+ enjoys being attracted to the (partially negative) oxygen end of the O–H bonds of water.

Finally, the $B^-(g)$ ions are surrounded by water molecules and enter solution:
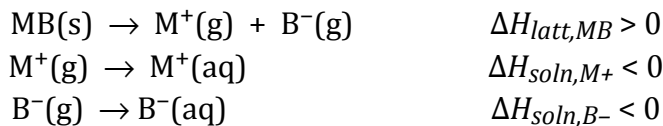
$$B^-(g) \rightarrow B^-(aq) \qquad \Delta H_{soln,B-} < 0$$

This quantity of energy is called the enthalpy of solution of the B^- ion $\Delta H_{soln,B-}$: it is a negative number because the energy of the system drops as the negative charge of B^- enjoys being attracted to the (partially positive) hydrogen end of the O–H bonds of water.
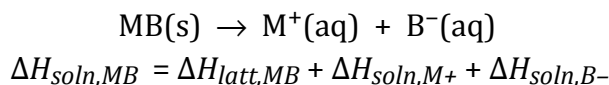
Figure 6-2 The thermodynamics of solution of the salt MB(s) in water for the case that $\Delta H_{soln,MB} < 0$, that is, dissolving MB(s) in water causes the temperature of the water to rise.
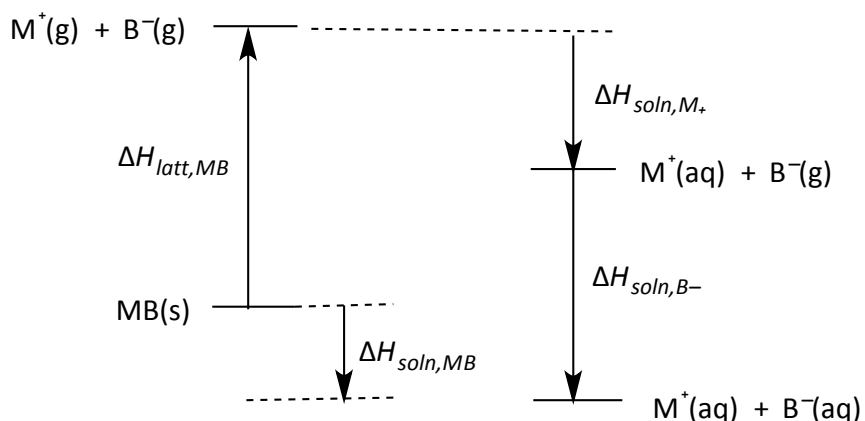
The enthalpy of solution $\Delta H_{soln,MB}$ of the salt MB(s) is thus given by

$$MB(s) \rightarrow M^+(g) + B^-(g) \qquad \Delta H_{latt,MB} > 0$$
$$M^+(g) \rightarrow M^+(aq) \qquad \Delta H_{soln,M+} < 0$$
$$B^-(g) \rightarrow B^-(aq) \qquad \Delta H_{soln,B-} < 0$$

Sum

$$MB(s) \rightarrow M^+(aq) + B^-(aq)$$
$$\Delta H_{soln,MB} = \Delta H_{latt,MB} + \Delta H_{soln,M+} + \Delta H_{soln,B-}$$

If the energy consumed in separating M^+ ions from B^- ions (i.e., $\Delta H_{latt,MB}$) is greater than the energy released when M^+ ions and B^- ions enter solution (i.e., $\Delta H_{soln,M+} + \Delta H_{soln,B-}$), then $\Delta H_{soln,MB} > 0$ and the overall solution process requires a net input of energy. In other words, dissolving MB in water causes the temperature of the water to fall.

Conversely, if the energy consumed in separating M^+ ions from B^- ions is less than the energy released when M^+ ions and B^- ions enter solution, then $\Delta H_{soln,MB} < 0$ and the overall solution process exhibits a net output of energy. In other words, dissolving MB in water causes the temperature of the water to rise.

Procedure

Determination of $\mathscr{C}_{cal}$

The first part of the lab session is devoted to evaluating $\mathscr{C}_{cal}$ of your calorimeter. You use an aluminum vessel as a calorimeter. The solutions that you are studying should be placed in the insulated inner cup of the calorimeter. The lid of the calorimeter is equipped with a cap (which should remain in place at all times), a thermometer, and a stirring ring (see Figure 6-1). It is essential that you use the same calorimeter for all your work.

Measure out about 50 mL of deionized water using a graduated cylinder. This volume corresponds to V_{cw} in Eqn. 6-2a; record the volume in your notebook. Pour the water into the inner cup of the calorimeter. Replace the lid and record the

temperature of the water at 30-sec intervals until it stabilizes. Be sure that the thermometer is suspended near, but not touching, the bottom or sides of the inner cup. Call the temperature at which the water stabilizes the initial temperature of the cold water $T_{cw,i}$; be sure to record it.

Using a Bunsen burner, bring about 125 mL of deionized water to a full boil in a beaker. Assume that the temperature of the boiling water is 100 °C; call this the initial temperature of the hot water $T_{hw,i}$. Please do not attempt to measure the temperature of the boiling water by dropping the thermometer into the beaker and leaving it there because the temperature sensor in the tip is damaged by prolonged contact with the extremely hot bottom of the beaker.

Put on heavy-duty gloves, extinguish the burner, pour the boiling water into the calorimeter, replace the lid, and vigorously agitate the contents of the calorimeter by moving the stirring ring up and down. Record the temperature 30 sec after the hot water is added to the calorimeter and at 30-sec intervals thereafter for at least 5 min; agitate the contents of the calorimeter throughout the data acquisition period. You want to observe a temperature decrease: if you don't, continue taking data past the recommended 5 min time period.

At the end of the data acquisition period, measure the total volume of water V_{tot} in the calorimeter using a graduated cylinder. The volume of hot water V_{hw} in Eqn. 6-2a equals the total volume of water in the calorimeter minus V_{cw}, that is, $V_{hw} = V_{tot} - V_{cw}$.

Working up the $\mathscr{C}_{cal}$ data
In your notebook prepare a plot of your $\mathscr{C}_{cal}$ data, plotting time along the x axis and temperature along the y axis. Call the time when you added the hot water to the cold water $t = 0$. The plot should look like Figure 6-2. For maximum accuracy the plot should take up a full page in your notebook. **Do not wait to prepare this plot and to calculate $\mathscr{C}_{cal}$ at home after lab!** If your $\mathscr{C}_{cal}$ data is no good, you must know so immediately so that you can take remedial action.

The value of $\mathscr{C}_{cal}$ is computed from Eq. 6-2a:

$$\mathscr{C}_{cal} = -\left(\frac{4.18\,\text{J}}{\text{mL}\cdot{}^\circ\text{C}}\right)\left[V_{hw}\left(\frac{\Delta T_{hw}}{\Delta T_{cw}}\right) + V_{cw}\right] \qquad \text{(Eqn. 6-2a)}$$

You already measured V_{cw} and V_{hw}. We now explain how the quantities ΔT_{cw} and ΔT_{hw} are determined. Let's define

$$\Delta T_{cw} = T_{cw,f} - T_{cw,i} \qquad \text{and} \qquad \Delta T_{hw} = T_{hw,f} - T_{hw,i}$$

$T_{cw,i}$ and $T_{hw,i}$ are the initial temperatures of the cold water and of the hot water, respectively; you already measured $T_{cw,i}$ and you are assuming that $T_{hw,i} = 100\ ^\circ\text{C}$. $T_{cw,f}$ and $T_{hw,f}$ are the final temperatures reached by the cold water and the hot water, respectively. It's clear that $T_{cw,f} = T_{hw,f}$ because the two bodies of water will have been mixed. Let's call this common final temperature of the mixture T_{mix}. Thus, the quantities we are looking for are

$$\Delta T_{cw} = T_{mix} - T_{cw,i} \qquad \text{and} \qquad \Delta T_{hw} = T_{mix} - T_{hw,i}$$

Use a ruler to draw a straight line through the data points lying on the downward sloping portion of the plot. Make sure that the line intercepts the vertical temperature axis erected at $t = 0$ (see Figure 6-3). The temperature T_{mix} corresponds to the y-intercept of the line and represents the temperature that the

Figure 6-3 Hypothetical data collected during the measurement of $\mathscr{C}_{cal}$. The hot water is added to the cold water in the calorimeter at $t = 0$. The line drawn through the data points intercepts the vertical
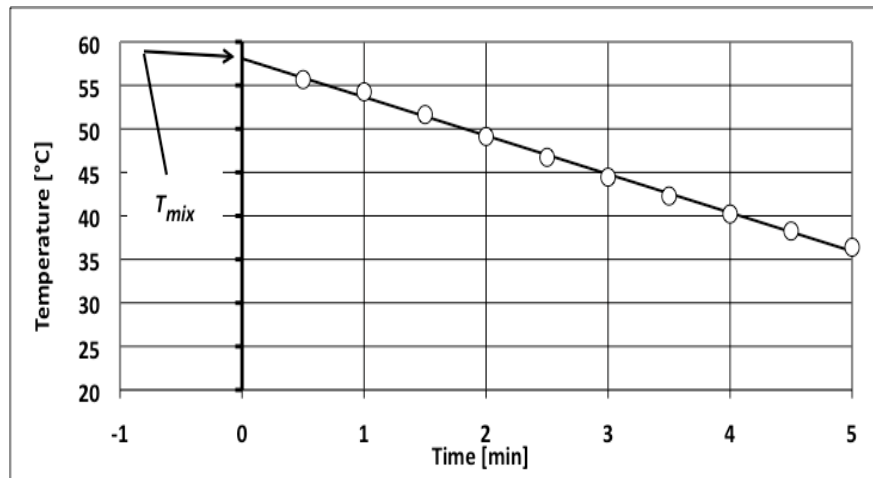
mixture would have exhibited if mixing and heat exchange were instantaneous.

We are at last ready to calculate $\mathscr{C}_{cal}$ using Eqn. 6-2b, which is a more detailed version of Eqn. 6-2a:

$$\mathscr{C}_{cal} = -\left(\frac{4.18\ \text{J}}{\text{mL}^\circ\ \text{C}}\right)\left[V_{hw}\left(\frac{T_{mix} - 100^\circ\text{C}}{T_{mix} - T_{cw,i}}\right) + V_{cw}\right]\quad\text{(Eqn. 6-2b)}$$

$\mathscr{C}_{cal}$ must be a positive number. If you arrive at a negative number, you did something wrong. A common mistake is boiling the water too far ahead of the time, leading to the addition of significantly less than 100 mL of boiling water. Check your calculations or repeat the experiment. If on the second try you once again arrive at a $\mathscr{C}_{cal}$ that is negative and you're sure that no errors have been committed, set $\mathscr{C}_{cal}$ equal to zero joules per degree Celsius and continue with the experiment.

Determination of $\Delta\mathscr{H}$ of NaCl(s) → Na⁺(aq) + Cl⁻(aq)

Weigh out 3–4 g of sodium chloride (NaCl(s)). In your notebook record the exact mass taken. Measure out with a graduated cylinder about 100 mL of deionized water and place it in the calorimeter; record the exact volume in your notebook. Cover the calorimeter with the lid and record the temperature of the water in the calorimeter at 30-sec intervals until it stabilizes. We will call the temperature at which the water stabilizes the initial temperature $T_{soln,i}$; be sure to record it.

Once the temperature of the water in the calorimeter has stabilized, add the NaCl(s) to the calorimeter, replace the lid, and agitate the contents of the calorimeter by moving the stirring ring up and down. Record the temperature 30 sec after the NaCl(s) is added to the calorimeter and at 30-sec intervals thereafter for at least 5 min; agitate the contents of the calorimeter throughout the data acquisition period. You must take data until you observe a temperature chnage of at least a few tenths of a degree; you may have to collect data past the recommended 5 min time period.

At the end of the data acquisition period, dispose of the solution in a hazardous waste receptacle. Rinse out the calorimeter and dispose of the rinses in a hazardous waste receptacle.

Repeat the procedure in this section, but this time weigh out 4–5 g of NaCl(s): you want data from two runs.

Determination of $\Delta\mathscr{H}$ of KCl(s) $\rightarrow$ K$^+$(aq) + Cl$^-$(aq)

Follow the instructions given in the previous section "Determination of $\Delta\mathscr{H}$ of NaCl(s) $\rightarrow$ Na$^+$(aq) + Cl$^-$(aq)", but now use 3–4 g of potassium chloride (KCl(s)) in the lower concentration run and 4–5 g of KCl(s) in the higher concentration run: you want data from two runs.

Determination of $\Delta\mathscr{H}$ of NH$_4$Cl(s) $\rightarrow$ NH$_4^+$(aq) + Cl$^-$(aq)

Follow the instructions given in the previous section "Determination of $\Delta\mathscr{H}$ of NaCl(s) $\rightarrow$ Na$^+$(aq) + Cl$^-$(aq)", but now use 3–4 g of ammonium chloride (NH$_4$Cl(s)) in the lower concentration run and 4–5 g of NH$_4$Cl(s) in the higher concentration run: you want data from two runs.

Data analysis

You use Eq. 6-5 to compute the molar enthalpy change $\Delta\mathscr{H}$ of solution:

$$\Delta\mathscr{H} = -\left(\frac{\mathcal{M}_{sys}\Delta T}{m_{sys}}\right)\left[\left(\frac{4.18\,\text{J}}{\text{mL}\cdot°\text{C}}\right)V_{soln} + \mathscr{C}_{cal}\right] \qquad \text{(Eqn. 6-5)}$$

where $\mathcal{M}_{sys}$ refers to the molar mass of the salt, m_{sys} refers to the mass in grams of salt used, V_{soln} refers to the volume of solution in the calorimeter (about 100 mL), and $\mathscr{C}_{cal}$ is the calorimeter constant you measured earlier.

The value of ΔT in Eqn. 6-5 is determined by the same technique employed in the measurement of $\mathscr{C}_{cal}$: time is plotted along the x axis and temperature along the y axis; call the time when you added the salt to the calorimeter $t = 0$; draw a straight line through the data points, making sure that the line intercepts the vertical temperature axis erected at $t = 0$. The temperature T_{mix} corresponds to the y-intercept of the line and represents the temperature that the mixture would have ex-

hibited if mixing and heat exchange were instantaneous. The quantity ΔT in Eqn. 6-5 is thus equal to $\Delta T = T_{mix} - T_{soln,i}$.

Name_____Lab Day_____Lab Time_____

Experiment 6 · Calorimetry

Lab report form **Page 1**

In (I.A), (II.A), (III.A), and (IV.A) you are asked to submit plots similar to Figure 6-3. Prepare a separate plot for each run. Give each plot a truly informative title (i.e., don't just call it "Run 1"), label the axes, and include appropriate units and divisions of those axes. Draw the straight line from which the value of T_{mix} is determined and write the value of T_{mix} on the plot. Do not submit small plots: use a whole sheet of paper. Scale the horizontal and vertical axes so that the data points occupy most of the area of the plot.

(I.A) Prepare a plot of the data collected in the measurement of $\mathscr{C}_{cal}$.

(I.B) Report the quantities needed for the calculation of $\mathscr{C}_{cal}$ according to Eqn. 6-2b:

$$\mathscr{C}_{cal} = -\left(\frac{4.18\ \text{J}}{\text{mL}\cdot°\ \text{C}}\right)\left[V_{hw}\left(\frac{T_{mix} - 100°\text{C}}{T_{mix} - T_{cw,i}}\right) + V_{cw}\right] \qquad \text{(Eqn. 6-2b)}$$

V_{hw} [mL]	V_{cw} [mL]	$T_{hw,i}$ [° C]	$T_{cw,i}$ [° C]	T_{mix} [° C]	T_{mix} – 100 °C [° C]	T_{mix} – $T_{cw,i}$ [° C]
		100 °C				

(I.C) Show the calculation of $\mathscr{C}_{cal}$ according to Eqn. 6-2b.

(I.D) What are the units of $\mathscr{C}_{cal}$?

(II.A) Prepare plots of the data collected in the measurement of $\Delta\mathscr{H}$ of NaCl(s) → Na⁺(aq) + Cl⁻(aq).

Experiment 6 · Calorimetry

Lab report form   **Page 2**

(II.B) Report the quantities needed in the calculation of $\Delta\mathcal{H}$ of NaCl(s) → Na$^+$(aq) + Cl$^-$(aq) according to Eqn. 6-5:

$$\Delta\mathcal{H} = -\left(\frac{\mathcal{M}_{sys}\Delta T}{m_{sys}}\right)\left[\left(\frac{4.18\,\text{J}}{\text{mL}\cdot°\text{C}}\right)V_{soln} + \mathcal{C}_{cal}\right] \qquad \text{(Eqn. 6-5)}$$

Calculate the value of $\Delta\mathcal{H}$ of each of your two runs and the mean value of $\Delta\mathcal{H}$. The variable $\mathcal{M}_{sys}$ refers to the molar mass of NaCl ($\mathcal{M}$ = 58.44 g/mol); the variable m_{sys} refers to the mass of NaCl used; V_{soln} refers to the volume of solution in the calorimeter.

Run	m_{sys} [g]	V_{soln} [mL]	$T_{soln,i}$ [° C]	T_{mix} [° C]	$\Delta T =$ $T_{mix} - T_{soln,i}$ [° C]	$\mathcal{C}_{cal}$	$\Delta\mathcal{H}$ [kJ/mol]
1							
2							
						mean	

(III.A) Prepare plots of the data collected in the measurement of $\Delta\mathcal{H}$ of KCl(s) → K$^+$(aq) + Cl$^-$(aq).

(III.B) Report the quantities needed in the calculation of $\Delta\mathcal{H}$ of KCl(s) → K$^+$(aq) + Cl$^-$(aq) according to Eqn. 6-5. Calculate the value of $\Delta\mathcal{H}$ of each of your two runs and the mean value of $\Delta\mathcal{H}$. The variable $\mathcal{M}_{sys}$ refers to the molar mass of KCl ($\mathcal{M}$ = 74.55 g/mol); the variable m_{sys} refers to the mass of KCl used; V_{soln} refers to the volume of solution in the calorimeter.

Run	m_{sys} [g]	V_{soln} [mL]	$T_{soln,i}$ [° C]	T_{mix} [° C]	$\Delta T =$ $T_{mix} - T_{soln,i}$ [° C]	$\mathcal{C}_{cal}$	$\Delta\mathcal{H}$ [kJ/mol]
1							
2							
						mean	

Experiment 6 · Calorimetry

Lab report form   **Page 3**

(IV.A) Prepare plots of the data collected in the measurement of $\Delta \mathscr{H}$ of $NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$.

(IV.B) Report the quantities needed in the calculation of $\Delta \mathscr{H}$ of $NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$ according to Eqn. 6-5. Calculate the value of $\Delta \mathscr{H}$ of each of your two runs and the mean value of $\Delta \mathscr{H}$. The variable $\mathscr{M}_{sys}$ refers to the molar mass of NH_4Cl ($\mathscr{M} = 53.49$ g/mol); the variable m_{sys} refers to the mass of NH_4Cl used; V_{soln} refers to the volume of solution in the calorimeter.

Run	m_{sys} [g]	V_{soln} [mL]	$T_{soln,i}$ [° C]	T_{mix} [° C]	$\Delta T =$ $T_{mix} - T_{soln,i}$ [° C]	$\mathscr{C}_{cal}$	$\Delta \mathscr{H}$ [kJ/mol]
1							
2							
						mean	

Post-lab questions

(1) In this experiment you measure enthalpy change $\Delta \mathscr{H}$. You are not measuring *standard* enthalpy change ΔH°. Why are the enthalpy changes you measure not standard? (Consulting your lecture textbook may help in answering this question.)

Name_____Lab Day_____Lab Time_____

Experiment 6 · Calorimetry

Lab report form   **Page 4**

(2.a) Look up the lattice energy $\Delta H_{latt,MB}$ of NaCl(s), KCl(s), and NH$_4$Cl(s) in units of kilo-joule per mole, indicating the source (URL or title of book) of the information. Some sources may denote the lattice energy by E or U and give the lattice energy as a negative or positive number; for our purposes lattice energies are all positive.

(2.b) Look up the ionic radius of Na$^+$, K$^+$, and NH$_4^+$ in units of meter on-line or in a book, indicating the source (URL or title of book) of the information. Many sources give ionic radii in units of angstrom (1 Å = 10^{-10} m).

Species	ΔH_{latt} [kJ/mol]	Source
NaCl(s)		
KCl(s)		
NH$_4$Cl(s)		
	Radius [m]	
Na$^+$		
K$^+$		
NH$_4^+$		

(2.c) Using the mean values of $\Delta\mathcal{H}$ reported in (II.B), (III.B), and (IV.B), the values of ΔH_{latt} in the table above, $\Delta H_{soln,Cl^-}$ = −352 kJ/mol, and the equation

$$\Delta\mathcal{H}_{soln,MCl} = \Delta H_{latt,MCl} + \Delta H_{soln,M+} + \Delta H_{soln,Cl-}$$

estimate the value of the enthalpy of solution $\Delta H_{soln,M+}$, where M$^+$ = Na$^+$, K$^+$, NH$_4^+$.

(2.d) Look up the enthalpy of solution of Na$^+$, K$^+$, NH$_4^+$, indicating the source (URL or title of book) of the information.

Species	My data $\Delta H_{soln,M+}$ [kJ/mol]	Looked up $\Delta H_{soln,M+}$ [kJ/mol]	Source
Na$^+$			
K$^+$			
NH$_4^+$			

Experiment 6 · Calorimetry

Lab report form   **Page 5**

(3.a) The lattice energy of the chloride salts studied here is strongly correlated to the ionic radius of the positive ion (Na^+, K^+, or NH_4^+); explain why.

(3.b) The enthalpies of solution of Na^+, K^+, and NH_4^+ are strongly correlated to their ionic radius; explain why.

(3.c) Draw a sketch showing how a water molecule orients itself as it approaches a K^+ ion.