

Your name: _____

Your instructor: _____

Lab Report

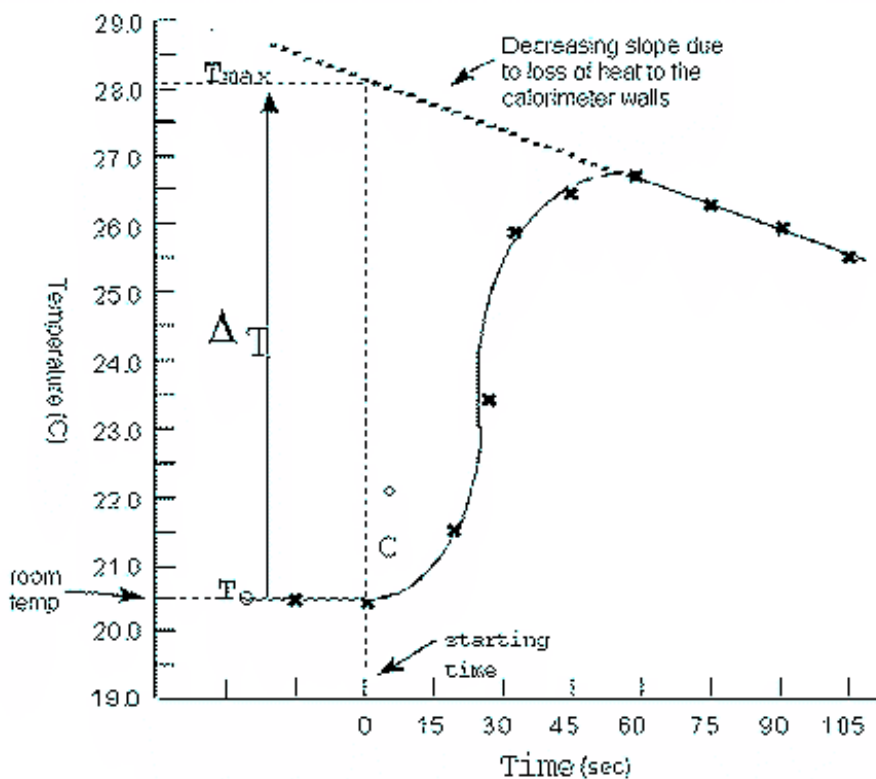
II. Calculations and Results

Show your calculations and results, NEATLY and with proper units, in the spaces provided. Enter final results in the answer spaces, with proper units and significant figures.

Accurate Determination of Final Temperature

For each run with the calorimeter, you need to know the initial and final temperatures in order to calculate Δt . The initial temperature is room temperature, t_{room} . The final temperature t_{final} is the maximum deviation from room temperature, after which the temperature begins drifting back towards room temperature. This drift is likely to begin between your timed measurements, which introduces error into your final temperature values. In addition, the reaction and heat release happens almost instantaneously, but there is a time delay for heating the water and calorimeter, and there is constant slight heat loss to the surroundings. To improve the precision of your final temperature measurements, make Excel graphs of each set of your time/temperature data, and extrapolate the line of temperature drift back to the time of mixing ($t = 0.0$ s), as shown in the graph below. The extrapolated value of temperature is your t_{final} for the run. Below is a sample graph and determination of t_{final} (labeled T_{max} on this graph). **If your calorimeter is tightly closed, the last part of your curves will be almost flat.**

Figure 1. Calorimetry: Determining Temperature Change by Extrapolating Cooling Curve



NOTE: You will need to graph data for **all** experiments in order to determine t_{final} accurately in each case. You need to prepare only one graph (from Part-1 data) to hand in with your report (p.7).

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Part 1. Heat Capacity of Calorimeter

The heat (q_{cal}) associated with a temperature change (Δt_{cal}) of the calorimeter depends on the heat capacity (C_{cal}) of the calorimeter, as follows:

$$q_{\text{cal}} = C_{\text{cal}} \cdot \Delta t_{\text{cal}} \quad (\text{Equation 1})$$

so the desired quantity C_{cal} equals $q_{\text{cal}}/\Delta t_{\text{cal}}$. The temperature change for the calorimeter is

$$\Delta t_{\text{cal}} = t_{\text{final}} - t_{\text{room}}$$

But how many joules (q_{cal}) of heat did the calorimeter gain? You added warm water to cool (room-temperature) water in the calorimeter, so the warm water heated both the cool water and the calorimeter. So the calorimeter gained the *difference* between the heat lost by the warm water, and the heat gained by the cool water:

$$q_{\text{cal}} = q_{\text{lost, warm water}} - q_{\text{gained, cool water}}$$

The heat (q) associated with the temperature change (Δt) of a certain mass (m) of water depends on the specific heat of water ($c = 4.184 \text{ J/g}\cdot\text{C}^\circ$), as follows:

$$q = m \cdot c \cdot \Delta t \quad (\text{Equation 2})$$

You can use this equation—twice—to calculate $q_{\text{lost, warm water}}$ and $q_{\text{gained, cool water}}$, in order to determine q_{cal} . Study the following sample calculation to see how.

Sample calculation of C_{cal} :

To determine the heat capacity of a calorimeter, a student added 50 mL warm water to a calorimeter containing 50 mL of room temperature water. The following temperatures were recorded:

- Temperature of 50.0 mL warm water: 38.2°C
- Temperature of 50.0 mL room temperature water: 21.2°C
- Temperature after mixing: 29.2°C
- Density of water: 1.00 g/mL, so 50.0 mL water equals 50.0 g water.

The heat lost by the warm water:

$$q = (50.0 \text{ g}) (4.184 \text{ J/g}\cdot\text{C}^\circ) (38.2 - 29.2^\circ\text{C}) = 1883 \text{ joules}$$

The heat gained by the cool water:

$$q = (50.0 \text{ g}) (4.184 \text{ J/g}\cdot\text{C}^\circ) (29.2 - 21.2^\circ\text{C}) = 1674 \text{ joules}$$

The heat lost to the calorimeter, q_{cal} , is the difference between the heat lost by the warm water and the heat gained by the room temperature water:

$$q_{\text{cal}} = q_{\text{lost, warm water}} - q_{\text{gained, cool water}} = 1883 - 1674 \text{ joules} = 209 \text{ joules}$$

The heat capacity of this calorimeter:

$$C_{\text{cal}} = q_{\text{cal}}/\Delta t_{\text{cal}} = 209 \text{ joules} / (29.2 - 21.2^\circ\text{C}) = \mathbf{26.1 \text{ joules} / ^\circ\text{C}}$$

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Follow the previous example to calculate C_{cal} for your calorimeter. Show your calculations here, emulating the example.

- Temperature of 50.0 mL warm water: _____ °C
- Temperature of 50.0 mL room temperature water: _____ °C
- Temperature of 100.0 mL water after mixing: _____ °C

The heat lost by the warm water:

$$q = \text{_____} = \text{_____} \text{ joules}$$

The heat gained by the cool water:

$$q = \text{_____} = \text{_____} \text{ joules}$$

The heat lost to the calorimeter, q_{cal} , is the difference between the heat lost by the warm water and the heat gained by the room temperature water:

$$q_{\text{cal}} = q_{\text{lost, warm water}} - q_{\text{gained, cool water}} = \text{_____} \text{ joules}$$

The heat capacity of your calorimeter:

$$C_{\text{cal}} = q_{\text{cal}} / \Delta t_{\text{cal}} = \text{_____} \text{ joules / } ^\circ\text{C}$$

Average C_{cal} = for your two trials

$$= \text{_____} \text{ joules / } ^\circ\text{C}$$

Consult the guidelines for graphs in this course. For **one** set of data in Part 1, Use Excel to prepare a publication-quality graph of calorimetry data. Use Figure 1, above, as a model. Use line-drawing and labeling tools (View:Toolbars:Drawing) in Excel to show the extrapolation -- **NO HAND WORK** allowed on graphs in this course. Adjust the size of your graph to fit the space below, and paste or tape the graph there. **Be sure that the graph fits all guidelines for graphs in this course, including title, axis labels, scaling, and so forth.**

**Paste graph
over this instruction.**

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Part 2. Heat of Neutralization: NaOH + HCl

When any process (in this case, a chemical reaction) occurs in the water of a calorimeter, the heat change for the process is transferred to both the water and the calorimeter:

$$q_{\text{process}} = -q_{\text{water}} - q_{\text{cal}} \quad (\text{Equation 3})$$

The value of q for the process is opposite in sign to the values for the water and calorimeter, because heat lost (or gained) by the process is gained (or lost) by the water and calorimeter. From the temperature change during an aqueous process in a calorimeter, you can calculate q_{cal} (Equation 1) and q_{water} (Equation 2), and thus you can calculate q_{process} . Because this process was carried out at constant pressure, the quantity q_{process} , when expressed in joules per mole of a specified reactant or product, is equal to the enthalpy change, ΔH for the process.

Study this sample calculation to see how to compute ΔH , in J/mole, for this reaction:

The following data were obtained for reaction of an acid and a base:

- Temperature of 50.0 mL of 2.0 M HCl before mixing: 23.0°C
- Temperature of 50.0 mL of 2.0 M NaOH before mixing: 23.0°C
- Temperature of 100 mL of solution after mixing: 28.0°C

Assume that the density (1 g/ml) and specific heat of the solution (4.184 J/g-°C) are the same as pure water.

To determine the heat gained by the solution, use Equation 2:

$$q_{\text{water}} = (100 \text{ g}) (4.184 \text{ J/g-}^\circ\text{C}) (28.0 - 23.0^\circ\text{C}) = 2.09 \times 10^3 \text{ J}$$

If the heat capacity of the calorimeter is 26.1 J/°C (from example above), then the heat gained by the calorimeter is given by Equation 1:

$$q_{\text{cal}} = (26.1 \text{ J/}^\circ\text{C}) (28.0 - 23.0^\circ\text{C}) = 131 \text{ J}$$

Use Equation 3 to calculate the heat change for the neutralization:

$$q_{\text{process}} = -(2092 \text{ J}) - (131 \text{ J}) = -2.22 \times 10^3 \text{ J}$$

After reaction, the final solution is 100.0 mL of 1.00 M NaCl, so it contains 0.100 moles of NaCl. So the heat of neutralization, in joules per mole of NaCl produced, is

$$\Delta H_{\text{neutralization}} = -2.22 \times 10^3 \text{ J} / 0.100 \text{ mol NaCl} = -2.22 \times 10^4 \text{ J/mol.}$$

NOTE: The temperature rose during this process, indicating an exothermic process, in agreement with $\Delta H < 0 \text{ J}$.

Follow the sample calculation to calculate ΔH for the reaction of NaOH and HCl. Show data, calculations, and results here, emulating the example:

- Temperature of 50.0 mL 2.0 M HCl before mixing: _____ °C
- Temperature of 50.0 mL 2.0 M NaOH before mixing: _____ °C
- Temperature of 100.0 mL of solution after mixing: _____ °C

$$q_{\text{water}} = \text{_____ J}$$

$$q_{\text{cal}} = \text{_____ J}$$

$$q_{\text{process}} = \text{_____ J}$$

$$\Delta H_{\text{neutralization}} = \text{_____ J/mol NaCl}$$

$$\text{Average } \Delta H_{\text{neutralization}} \text{ for your two trials} = \text{_____ J/mol NaCl}$$

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Part 3. Heat of Solution of NH_4Cl

Following the procedure of the sample calculation in Part 2, calculate ΔH (J/mol), the heat of solution for ammonium chloride. Enter data, calculations, and results here, emulating the example.

Express $\Delta H_{\text{solution}}$ as J/mol NH_4Cl dissolved.

• Temperature of 100.0 mL of water before adding NH_4Cl : _____ °C

• Temperature after mixing: _____ °C

$$q_{\text{water}} = \text{_____} = \text{_____} \text{ J}$$

$$q_{\text{cal}} = \text{_____} = \text{_____} \text{ J}$$

$$q_{\text{process}} = \text{_____} = \text{_____} \text{ J}$$

$$\text{moles } \text{NH}_4\text{Cl} \text{ dissolved} = \text{_____} = \text{_____} \text{ mol}$$

$$\Delta H_{\text{solution}} = \text{_____} = \text{_____} \text{ J/mol } \text{NH}_4\text{Cl}$$

$$\text{Average } \Delta H_{\text{solution}} \text{ for your two trials} = \text{_____} \text{ J/mol } \text{NH}_4\text{Cl}$$

Part 4. Heat of Reaction: $\text{Ca (s)} + \text{HCl (aq)}$

Following the procedure of the sample calculation in Part 2, calculate the molar heat of reaction, ΔH (J/mol), for the the reaction of calcium with hydrochloric acid. Enter data, calculations, and results here, emulating the example. Express $\Delta H_{\text{solution}}$ as J/mol Ca reacted.

• Temperature of 100.0 mL of 2.0 M HCl before adding Ca: _____ °C

• Temperature after mixing: _____ °C

$$q_{\text{water}} = \text{_____} = \text{_____} \text{ J}$$

$$q_{\text{cal}} = \text{_____} = \text{_____} \text{ J}$$

$$q_{\text{process}} = \text{_____} = \text{_____} \text{ J}$$

$$\text{moles Ca reacted} = \text{_____} = \text{_____} \text{ mol}$$

$$\Delta H_{\text{reaction}} = \text{_____} = \text{_____} \text{ J/mol Ca}$$

$$\text{Average } \Delta H_{\text{reaction}} \text{ for your two trials} = \text{_____} \text{ J/mol Ca}$$

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III. Questions

1. Write full balanced chemical equations, specifying states (s, l, g, aq) of each component, for the reactions of Parts 2, 3, and 4. Then write net ionic equations for each. Answer for all Parts, even if you did not do all parts.

Part 2:

Chemical equation:

Net ionic equation:

Part 3:

Chemical equation:

Net ionic equation:

Part 4:

Chemical equation:

Net ionic equation:

2. Characterize the reactions you studied as endothermic or exothermic. Explain briefly. Answer only for the parts you carried out in lab.

Part 2 ____: Circle one: endothermic exothermic neither

Explain:

Part 3 ____: Circle one: endothermic exothermic neither

Explain:

Part 4 ____: Circle one: endothermic exothermic neither

Explain:

3. For an **endothermic** reaction in a calorimeter, if t_{final} were determined by the maximum deviation of the temperature reading from room temperature (without extrapolation to time zero), the calculated value of $\Delta H_{\text{reaction}}$ would be be *higher lower same* (circle one), in comparison to the **actual** value. Explain, by drawing a graph of the expected data, the effect of *not* extrapolating to get t_{final} , and the resulting effect on calculated $\Delta H_{\text{reaction}}$. Use the back of this page if necessary.

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4. a) For the reaction(s) that you studied (parts 2, 3, and/or 4), calculate the standard enthalpy of reaction (ΔH°), using standard heats of formation (ΔH°_f) from the appendix of your classroom text book. Express the answer in kJ/mol.

b) **Then** compare the standard enthalpy you calculated with the enthalpy change you found experimentally (expressed in kJ/mol). If the difference is more than 10%, suggest possible reasons, taking into consideration the limitations of your lab apparatus (such as exchange of heat with surroundings, or precision of determining temperature changes). Make sure that your explanation fits the discrepancy.

Report

At the time and date specified on your lab schedule, hand in this completed form.

REMINDER: Even though you worked in pairs, your report should be your own individual work.