SI Chemistry - Full Discipline Demo

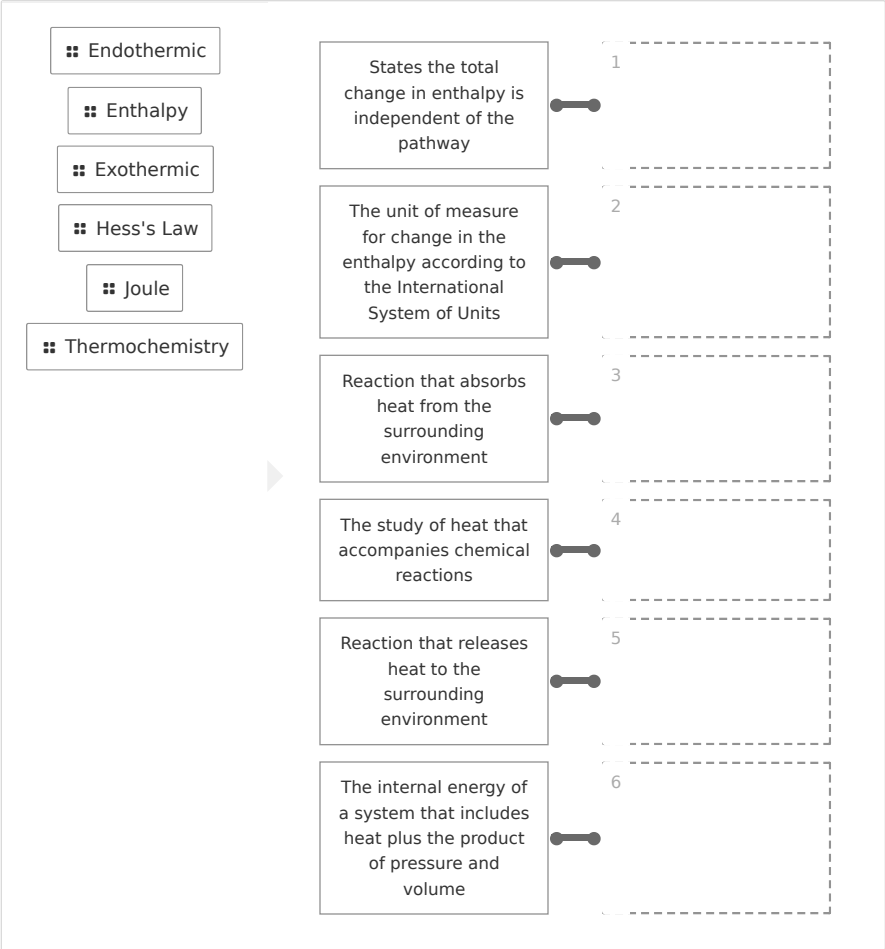
Hess's Law

Final Report - Answer Guide

Institution	Science Interactive University
Session	SI Chemistry - Full Discipline Demo
Course	SI Chemistry - Full Discipline Demo
Instructor	Sales SI Demo

Test Your Knowledge

Match each term with the best description.

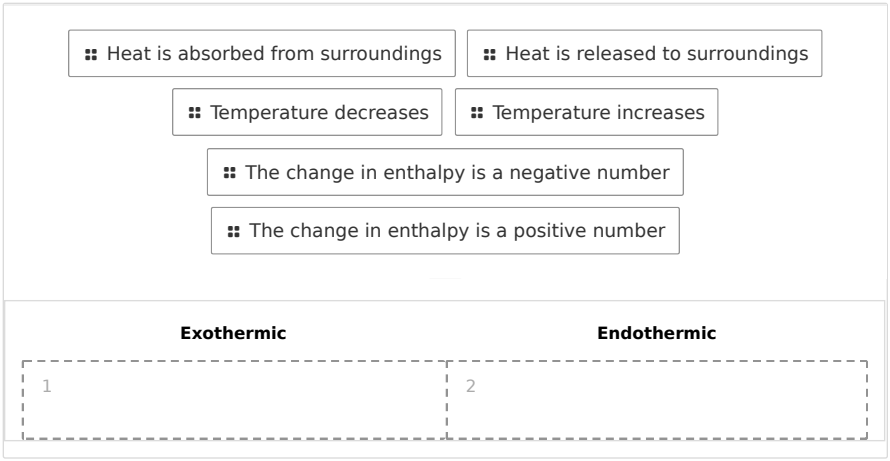


Correct answers:

1 Hess's Law 2 Joule 3 Endothermic 4 Thermochemistry

5 Exothermic 6 Enthalpy

Categorize each statement as representing an exothermic or endothermic reaction.



Correct answers:

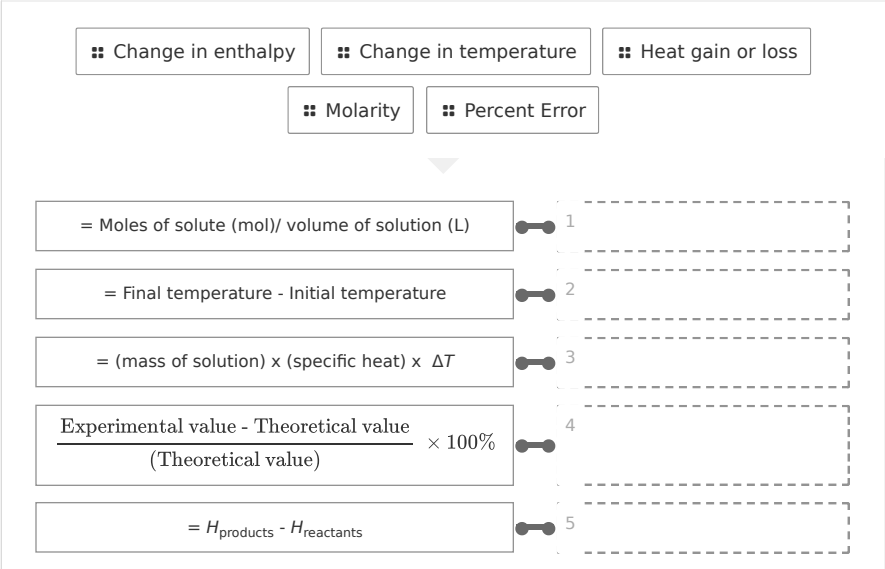
1 Heat is released to surroundings Temperature increases

The change in enthalpy is a negative number

2 Heat is absorbed from surroundings Temperature decreases

The change in enthalpy is a positive number

Match term with the correct equation.



Correct answers:

1 Molarity 2 Change in temperature 3 Heat gain or loss

4 Percent Error 5 Change in enthalpy

Exploration

When an _____ reaction occurs in an aqueous solution, heat will be released to the water.

- endothermic
- exothermic ✔

Enthalpy is the external energy of the system, *E*, which includes heat plus the product of pressure and volume.

- True
- False ✔

The change in enthalpy of a reaction is equal to the enthalpy of the reactants minus the enthalpy of the products.

- True
- False ✔

When a reaction is exothermic, the enthalpy of the products is _____ the reactants.

- greater than
- less than ✔
- equal to

Hess's law states that the total change in enthalpy is independent of the pathway.

- True ✔
- False

Exercise 1

What were the maximum temperatures for each of the reactions? Use the scatter plots in Graph 1, Graph 2, and Graph 3 to support your answers.

The maximum temperatures for each of the reactions were 37.4, 25.5, and 35.2 respectively.

Calculate the heat loss or heat gain of the 3 solution mixtures (q_{rxn}). Then calculate the ΔH for each reaction. Write the calculations on a sheet of paper showing all of your work. Take a photo of your work for each reaction and upload the images into Photo 1, Photo 2, and Photo 3.

First Reaction:

$$\mathrm{NaOH(aq)} + \mathrm{HCl(aq)} \rightarrow \mathrm{H_2O(l)} +$$

$$2\mathrm{M} = \frac{2\ \text{moles}}{\cancel{1\ \text{Liter}}} \times \frac{\cancel{1\ \text{Liter}}}{1000\ \cancel{\text{mL}}} \times 10\ \cancel{\text{mL}} =$$

$$\text{Density} = \frac{1.02\ \text{g}}{\cancel{1\ \text{mL}}} \times 20\ \cancel{\text{mL}} = 20.4\ \text{g} =$$

$$q_{\text{rxn}} = (20.4\ \text{g}) \times (4.184\text{J/g}\,^{\circ}\text{C}) \times (-13.4^{\circ}$$

$$\Delta H = -1.14\ \text{kJ}/0.02\ \text{mol} = -57\ \text{kJ} /$$

Second Reaction:

$$NaOH(aq) + NH_4Cl(aq) \rightarrow NH_3(g$$

$$2M = \frac{2 \text{ moles}}{1 \cancel{\text{Liter}}} \times \frac{1 \cancel{\text{Liter}}}{1000 \cancel{\text{mL}}} \times 10 \cancel{\text{mL}} =$$

$$\text{Density} = \frac{1.02 \text{ g}}{1 \cancel{\text{mL}}} \times 20 \cancel{\text{mL}} = 20.4 \text{ g} =$$

$$q_{rxn} = (20.4 \text{ g}) \times (4.184 \text{J/g}^\circ\text{C}) \times (-0.7^\circ$$

$$\Delta H = -0.060 \text{ kJ}/0.02 \text{ mol}$$

Third Reaction:

$$NH_3(aq) + HCl(aq) \rightarrow N$$

$$2M = \frac{2 \text{ moles}}{1 \cancel{\text{ Liter}}} \times \frac{1 \cancel{\text{ Liter}}}{1000 \cancel{\text{ mL}}} \times 10 \cancel{\text{ mL}} =$$

$$\text{Density} = \frac{1.02 \text{ g}}{1 \cancel{\text{ mL}}} \times 20 \cancel{\text{ mL}} = 20.4 \text{ g} =$$

$$q_{rxn} = (20.4 \text{ g}) \times (4.184 \text{J/g}\,^\circ\text{C}) \times (-10.2$$

$$\Delta H = -0.875 \text{ kJ}/0.02 \text{ mol} =$$

Use Hess's law and your determination of Δ*H* for the first 2 reactions to determine Δ*H* for the third reaction:

NH_3 + HCl → NH_4Cl.

Because the second reaction is reversed to get the third, the sign from the second reaction will change to a positive value instead of a negative value (see the background section for addition of the reactions). Therefore, you sum (-57 kJ) + (3 kJ) = -54 kJ.

Compare the results of question 3 with the experimental results of the reaction:

NH_3 + HCl → NH_4Cl (calculate the percent error)

Percent Error = [(-54 kJ / mol) - (-43.75 kJ / mol)] / (-43.75 kJ / mol) x 100% = 30.3%

Use the thermodynamic quantities given below to calculate the theoretical Δ*H* for this reaction:

NH_3 + HCl → NH_4Cl

- **$\Delta H°_f$ for NH_3 (aq) = - 80.29 kJ/mol**
- **$\Delta H°_f$ for HCl (aq) = - 167.2 kJ/mol**
- **$\Delta H°_f$ for NH_4^+ (aq) = - 132.5 kJ/mol**
- **$\Delta H°_f$ for Cl^- (aq) = - 167.2 kJ/mol**

The theoretical ΔH = -52.2 kJ

What is the percent error of experimental results when using the theoretical value calculated in Question 5?

Percent Error = [(-43.75 kJ / mol) - (-52.2 kJ / mol)] / (-52.2 kJ / mol) x 100% = 16.2%

Were the reactions performed in this laboratory exercise exothermic or endothermic? How could you determine this?

The reactions were exothermic because they released heat (the temperature rose).

Define Hess's law. Did your experimental results support Hess's law? Use your data to explain your answer.

Hess's Law states that if a reaction is the sum of two or more other reactions, the Δ*H* for the overall process must be the sum of the Δ*H* values of the constituent reactions. The thermal energy absorbed or produced by a chemical process reflects a difference between the enthalpy, the reactants and products (Δ*H*).

Student answers may vary somewhat when using data to support or refute Hess's Law. Students should find that the enthalpies of the intermediate reactions were comparable to that of the system as a whole. HOL found a 30.3% error in the experimental values.

What are some possible sources of error in this experiment?

Possible sources of error include heat escaping from the calorimeter, the thermometer not capturing the temperature of the liquid correctly, the temperature of the ambient air, or not getting the exact temperature when first mixing the two chemicals.

Data Table 1: Temperature (°C) of the Solution (NaOH + HCl)
(SAMPLE ANSWER BELOW)

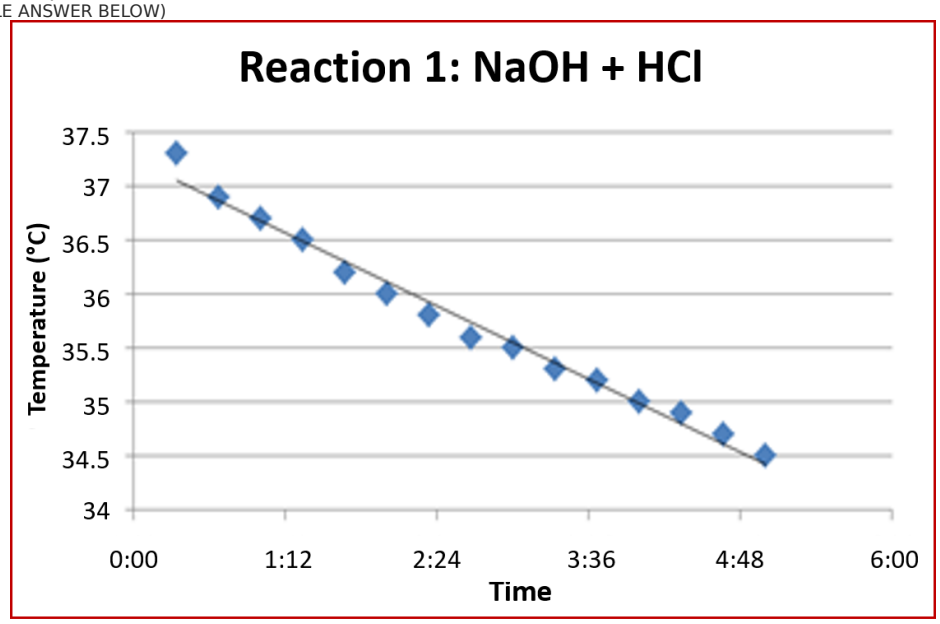
Initial Temperature of NaOH (°C)	23.4
Initial Temperature of HCl (°C)	24.4
Average Initial Temperature of Two Separate Solutions (°C)	23.9
Highest Temperature of Mixture (May Be Extrapolated from Graph) (°C)	37.3
Change in Temperature of Mixture (°C), Δ*T*	13.4

Data Table 2: Temperature of Solution Every 20 Seconds
(SAMPLE ANSWER BELOW)

Time after mixing (min:sec)	Temperature (°C)
Initial	37.3
0:20	37.3
0:40	36.9
1:00	36.7
1:20	36.5
1:40	36.2
2:00	36.0
2:20	35.8
2:40	35.6

3:00	35.5
3:20	35.3
3:40	35.2
4:00	35.0
4:20	34.9
4:40	34.7
5:00	34.5

Graph 1: Reaction 1 (NaOH + HCl)
(SAMPLE ANSWER BELOW)



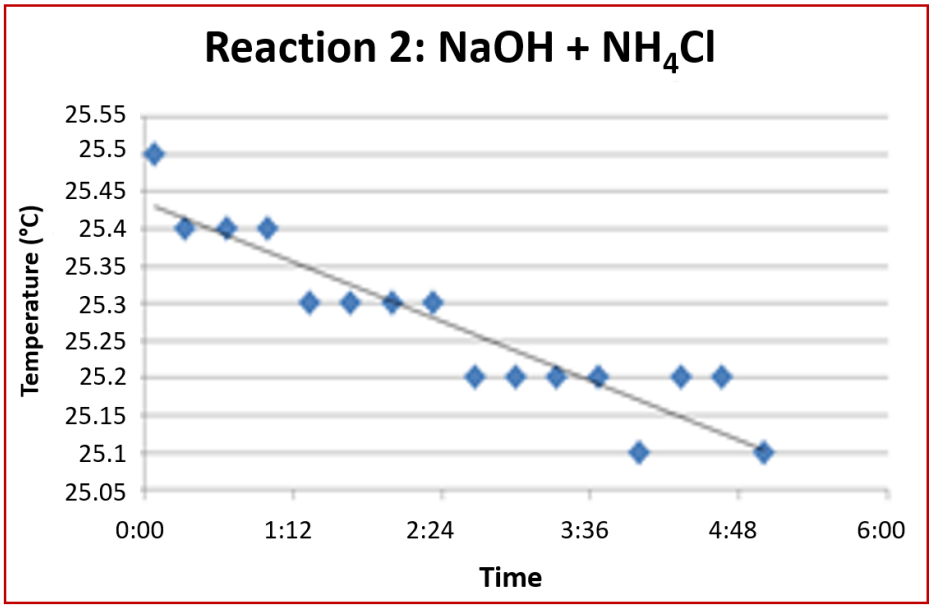
Data Table 3: Temperature (°C) of the Solution (NaOH + NH4Cl)
(SAMPLE ANSWER BELOW)

Initial Temperature of NaOH (°C)	25.4
Initial Temperature NH_4Cl (°C)	24.2
Average Initial Temperature (°C)	24.8
Highest Temperature of Mixture (May Be Extrapolated from Graph) (°C)	25.5
Change in Temperature of Mixture (°C), ΔT	0.7

Data Table 4: Temperature of Solution Every 20 Seconds
(SAMPLE ANSWER BELOW)

Time after mixing (min:sec)	Temperature (°C)
Initial	25.5
0:20	25.4
0:40	25.4
1:00	25.4
1:20	25.3
1:40	25.3
2:00	25.3
2:20	25.3
2:40	25.2
3:00	25.2
3:20	25.2
3:40	25.2
4:00	25.1
4:20	25.2
4:40	25.5
5:00	25.1

Graph 2: Reaction 2 (NaOH + NH4Cl)
(SAMPLE ANSWER BELOW)



Data Table 5: Temperature (°C) of the Solution (NH3 + HCl)
(SAMPLE ANSWER BELOW)

Initial Temperature of NH_3 (°C)	24.9
Initial Temperature HCl (°C)	25.0
Average Initial Temperature (°C)	24.95
Highest Temperature of Mixture (May Be Extrapolated from Graph)(°C)	35.2
Change in Temperature of Mixture (°C), ΔT	10.25

Data Table 6: Temperature of Solution Every 20 Seconds
(SAMPLE ANSWER BELOW)

Time after mixing (min:sec)	Temperature (°C)
Initial	35.2
0:20	34.3
0:40	34.0
1:00	33.8
1:20	33.6
1:40	33.4
2:00	33.2
2:20	33.2
2:40	33.1
3:00	33.0
3:20	32.8
3:40	32.8
4:00	32.8
4:20	32.6

4:40	32.5
5:00	32.4

Graph 3: Reaction 3 (NH3 + HCl)
(SAMPLE ANSWER BELOW)

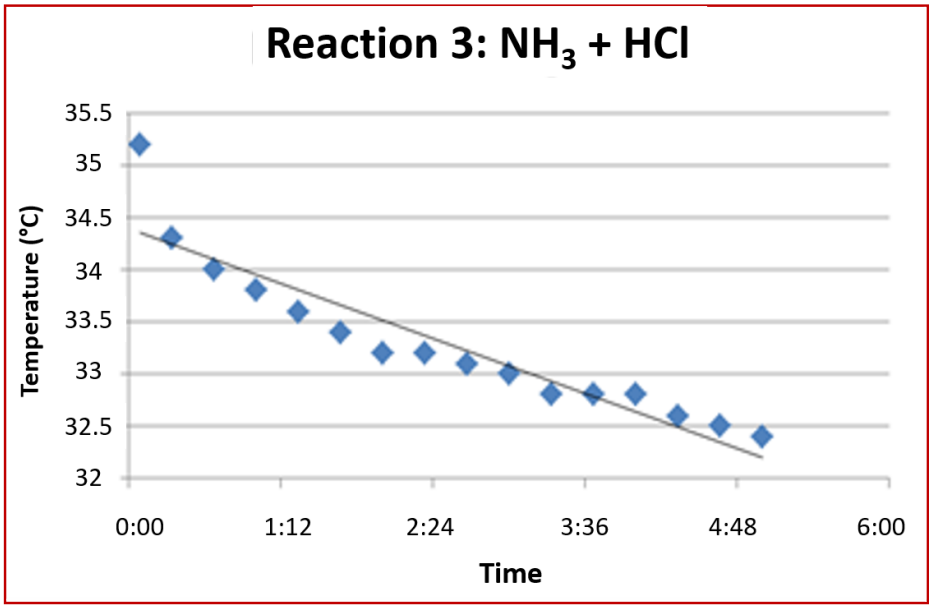


Photo 1: Reaction 1: NaOH + HCl Calculations
(SAMPLE ANSWER BELOW)

$$NaOH(aq) + HCl(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

$$2M = \frac{2 \text{ moles}}{1 \text{ Liter}} \times \frac{1 \text{ Liter}}{1000 \text{ mL}} \times 10 \text{ mL} = 0.02 \text{ moles}$$

Density = $\frac{1.02\ g}{1\ mL}$ x 20 mL = 20.4 g = total mass

qrxn = (20.4 g) x (4.184 J / g°C) x (-13.4°C) = -1143.74 J = -1.14 kJ

ΔH = -1.14 kJ / 0.02 mol = -57 kJ / mol

Photo 2: Reaction 2: NaOH + NH4Cl Calculations
(SAMPLE ANSWER BELOW)

$NaOH(aq) + NH_4Cl(aq) \rightarrow NH_3(g) + NaCl(aq) + H_2O(l)$

2M = $\frac{2\ moles}{1\ Liter}$ x $\frac{1\ Liter}{1000\ mL}$ x 10 mL = 0.02 moles

Density = $\frac{1.02\ g}{1\ mL}$ x 20 mL = 20.4 g = total mass

qrxn = (20.4 g) x (4.184 J / g°C) x (-0.7°C) = -59.74J = -0.060 kJ

ΔH = -0.060 kJ / 0.02 mol = -3 kJ / mol

Photo 3: Reaction 3: NH3 + HCl Calculations
(SAMPLE ANSWER BELOW)

$NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$

$$2\ M = \frac{2\ moles}{1\ Liter} \times \frac{1\ Liter}{1000\ mL} \times 10\ mL = 0.02\ moles$$

$$Density = \frac{1.02\ g}{1\ mL} \times 20\ mL = 20.4\ g = total\ mass$$

$qrxn = (20.4\ g) \times (4.184\ J/g^{\bullet}C) \times (-10.25^{\bullet}C) = -874.9\ J = -0.875\ kJ$

$\Delta H = -0.875\ kJ / 0.02\ mol = -43.75\ kJ/mol$

Competency Review

In an endothermic reaction _____.

- heat is released and temperature increases
- heat is released and temperature decreases
- heat is absorbed and temperature decreases ✔
- heat is absorbed and temperature increases

A chemical reaction that releases heat to the surrounding environment is considered _____.

- endothermic
- exothermic ✔

The amount of heat given off or absorbed by a chemical reaction is called _____.

- an exothermic reaction
- the change in enthalpy ✔
- a Joule
- the product of pressure and volume

Enthalpy is denoted by the _____ symbol.

- *H* ✔
- J
- Δ*E*
- F

The change in enthalpy of a reaction is equal to the enthalpy of the _____ minus the enthalpy of the _____.

- products; reactants ✔
- reactants; products
- None of the above

When an _____ reaction occurs the products have a greater enthalpy than the reactants, whereas during an ___ reaction the enthalpy of the products is less than the enthalpy of the reactants.

- exothermic; endothermic
- endothermic; exothermic ✔

Hess's Law indicates that if a reaction takes place in one or multiple steps, the change in enthalpy for the overall process must be the sum of the change in enthalpy of the constituent reactions.

- True ✔
- False

Calculate the percent error of a chemical reaction with a theoretical value of -45.68 kJ/mol and an experimental value of -54.78 kJ/mol.

- 20.10%
- 16.61%
- 19.92% ✔
- 21.37%

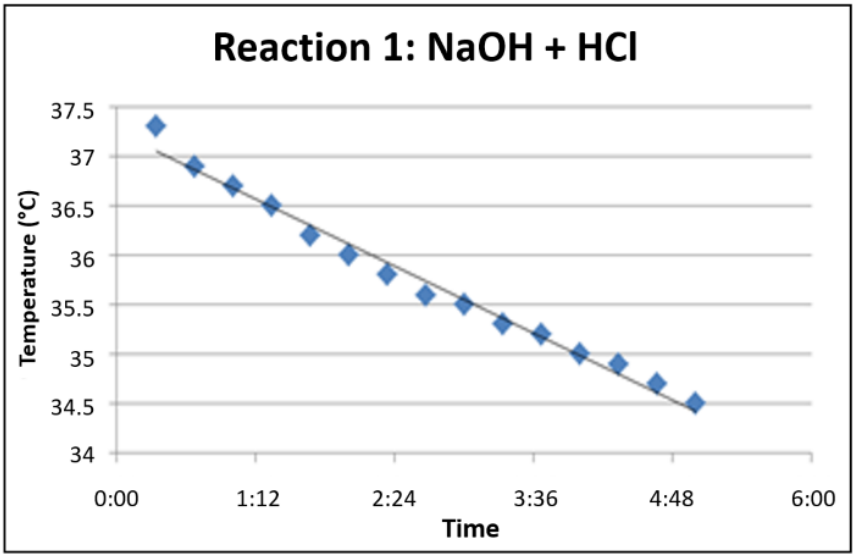
What is the ΔH of a reaction where the mass of the solution is 18.7 g at a specific heat of 5.008 J/g°C and the system experiences a change in temperature of 7.2°C?

- 674.27 J
- -674.27 J ✔
- -67.42 J
- 67.42 J

If an experiment has a q_{rxn} of 518.07 J, calculate the change in enthalpy of the reaction if there were 0.04 mol of product in the solution.

- 12.95 kJ/mol ✔
- -12951.75 kJ/mol
- -12.95 kJ/mol
- 7.72 kJ/mol

In an experiment, NaOH and HCl are mixed. Based on the graph, what is the change in temperature of the mixture if the initial temperature of the NaOH was 23.0°C and the initial temperature of the HCl was 25.2°C?



- 1.3°C
- 12.7°C
- 13.2°C ✔
- 10.4°C

Extension Questions

Worldwide, scientists are working on different options for clean energy to replace the burning of fossil fuels. One option that could potentially be successful is the use of hydrogen gas. This gas can be generated through a reaction between carbon and water.

$$C(s) + 2H_2O(g) \longrightarrow CO_2(g) + 2H_2(g)\ \Delta H = ?$$

1) $CO_2(g) \longrightarrow C(s) + O_2(g)\ \ \Delta H = 393.5\ KJ$

2) $2H_2O(g) \longrightarrow 2H_2(g) + O_2(g)\ \ \Delta H = 483.6\ KJ$

Based on what you have learned in this lab, use the chemical equation and Hess's law to determine the heat of formation and if this reaction is exothermic or endothermic.

(SAMPLE ANSWER BELOW)

Equation 1 must be changed to C(s) + O2 (g) --> CO2 (g) H1 = -393.5 KJ

Equation 2 is in the correct order.

2H2O --> 2H2(g) + O2(g) H2 = 483.6 KJ

Adding the two equations together cancels out the oxygen in equation 1 and 2 to give the desired reaction equation.

H1 + H2 = -393.5 KJ + 483.6 KJ = 90.1 KJ

This means the heat of formation for this reaction gives a positive enthalpy, which makes this reaction endothermic.