

## Experiment 2 – Enthalpies of Various Reactions

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# Enthalpy of Various Reactions

## Introduction

Enthalpy is the amount of heat that is absorbed or released by a system when the pressure is kept constant. Enthalpy is typically expressed as a change of enthalpy, where heat is lost or gained during a chemical reaction. The change in enthalpy is related to the internal energy and the first law of thermodynamics, such that no heat is created nor destroyed in an isolated system, but rather transformed from one form to another.

An exothermic reaction occurs when the system releases heat to its surroundings. This results in a negative heat value for the system, signifying that the system's temperature decreased while the surroundings' temperature increased. An endothermic reaction occurs when the system absorbs heat from its surroundings. This results in a positive heat value for the system, indicating that the system's temperature increased while the surroundings' temperature decreased.

Calorimetry is the science of measuring heat. The heat of a reaction is measured using a calorimeter: a device that can be as simple as a styrofoam cup with a lid, which keeps the contained solution relatively isolated from the environment. When a reaction occurs within the calorimeter, the change in temperature is recorded and used to measure the heat lost or gained during the reaction.

The enthalpy of a metal can be measured by determining the specific heat capacity of the metal. The specific heat capacity is the amount of heat required to change the temperature of one gram of a substance by one degree Celsius or one Kelvin. In this experiment, a range of equations are used to determine the experimental specific heat capacity of a metal, its molar mass, and the error of these findings in relation to their actual values:

1)  $\Delta T_{\text{water}} = T_f - T_i$

This equation is used to determine the temperature change of the water, the temperature change of the metal, or the temperature change of any other solution. The change in temperature is the difference between the final temperature and the initial temperature.

2)  $q = mc\Delta T$

This equation is used to determine the heat absorbed by the water by multiplying the mass of the water by its specific heat capacity and its change in temperature.

3)  $-q(\text{metal}) = q(\text{water})$

This equation is used to determine the heat released by the metal, following the principle that the heat lost by the system is equal to the heat gained by the surroundings.

4)  $C_{\text{metal}} = -m_{\text{water}}C_{\text{water}}\Delta T_{\text{water}} / m_{\text{metal}}\Delta T_{\text{metal}}$

This equation is used to determine the experimental specific heat capacity of the metal by dividing the negative value of the heat absorbed by the water by the mass of the metal multiplied by its temperature change.

$$5) \% \text{ error} = (|C_{\text{exp}} - C_{\text{actual}}|/C_{\text{actual}}) \times 100\% \quad \& \quad \% \text{ error} = (|MM_{\text{exp}} - MM_{\text{actual}}|/MM_{\text{actual}}) \times 100\%$$

These types of equations are used in each part of the experiment to determine the percent error in the experiment. By finding the difference between an experimental value and an actual value, then dividing it by the actual value and multiplying it by 100, the percent error can be determined thus allowing for the evaluation of the success of the experiment.

The enthalpy of neutralization is the change in enthalpy when an acid and a base undergo a neutralization reaction to produce water and salt. In this experiment, a variety of equations are used to calculate the enthalpy of neutralization of a base using different acids:

$$1) V_{\text{solution}} = V_{\text{Acid}} + V_{\text{Base}}$$

This equation is used to determine the total volume of the solution. This is done by adding the volumes of each individual component in the solution, in this case the acid and the base.

$$2) m_{\text{solution}} = V_{\text{solution}} \times d_{\text{solution}} \qquad m_{\text{solution}} = m_{\text{calorimeter+solution}} - m_{\text{calorimeter}}$$

These equations are used to determine the mass of the solution by either multiplying the total volume of the solution by the density of the solution, or simply by using a balance to find the actual mass of the solution.

$$3) q_N = -m_{\text{total}}C_{\text{solution}}\Delta T_{\text{solution}}$$

This equation is used to determine the heat of neutralization, which is a negative value, indicating that the heat is released by the solution. This value is calculated by multiplying the total mass of the solution by the specific heat capacity of the solution and the change in temperature of the solution.

$$4) n_{\text{OH}^-} = V_{\text{NaOH}} \times C_{\text{NaOH}} \times (1 \text{ mol OH}^-/1 \text{ mol NaOH}) \quad \& \quad n_{\text{OH}^-} = n_{\text{H}_2\text{O}}$$

These equations are used to determine the number of moles of OH<sup>-</sup> and H<sub>2</sub>O. The ratio of OH<sup>-</sup> to NaOH is 1:1 as is H<sub>2</sub>O, and so both number of moles can be determined by multiplying the volume of NaOH by the concentration of NaOH.

$$5) \Delta_N H^\circ = q_N / n_{\text{H}_2\text{O}}$$

This equation is used to determine the enthalpy of neutralization per mole of water by dividing the heat of neutralization by the number of moles of water.

The enthalpy of solution or enthalpy of dissolution is the enthalpy associated with the dissolution of a solute in a solvent under constant pressure conditions. In this experiment, many equations are used to determine the enthalpy of solution of an unknown salt.

$$1) q_{\text{solution}} = (-q_{\text{water}}) (-q_{\text{salt}}) \\ q_{\text{solution}} = (-m_{\text{water}}C_{\text{water}}\Delta T_{\text{water}}) + (-m_{\text{salt}}C_{\text{salt}}\Delta T_{\text{salt}})$$

This equation is used to determine the heat of solution by determining the heat of water and the heat of the salt and adding the two values.

2)  $\Delta_s H^\circ = q_{\text{solution}} / n_{\text{salt}}$

This equation is used to determine the enthalpy of solution, by dividing the heat of solution by the number of moles of salt. The number of moles of salt is determined by dividing the mass of salt by the molar mass of the salt.

*References:*

What in the World ISN'T Chemistry, Dr. R. Venkateswaran, Experiment 2, pp. 28-31

[http://chemwiki.ucdavis.edu/Physical\\_Chemistry/Thermodynamics/State\\_Functions/Enthalpy](http://chemwiki.ucdavis.edu/Physical_Chemistry/Thermodynamics/State_Functions/Enthalpy), UC Davis, Enthalpy

[http://chemwiki.ucdavis.edu/Physical\\_Chemistry/Equilibria/Le\\_Chatelier's\\_Principle/Effect\\_Of\\_Temperature\\_On\\_Equilibrium\\_Composition/Exothermic\\_Versus\\_Endothermic\\_And\\_K](http://chemwiki.ucdavis.edu/Physical_Chemistry/Equilibria/Le_Chatelier's_Principle/Effect_Of_Temperature_On_Equilibrium_Composition/Exothermic_Versus_Endothermic_And_K), UC Davis, Exothermic vs. Endothermic and K

<http://www.science.uwaterloo.ca/~cchieh/cact/c120/calorimetry.html>, University of Waterloo, Calorimetry

**Procedure:**

As described in the lab manual: What in the World ISN'T Chemistry, Dr. R. Venkateswaran, Experiment 2, pp.32-36

## **Data and Observation/Results:**

Table 1. Part 1 of the experiment – enthalpy of a metal

Data	Trial 1	Trial 2
Identity of the metal	Copper	Copper
Mass of the metal (g)	13.24	13.49
Volume of distilled water (mL)	20.0	20.0
Mass of the empty calorimeter cup (g)	1.8972	1.8972
Mass of the calorimeter cup with water (g)	21.1836	21.1070
Mass of water (g)	19.2864	19.2098
Temperature of boiling water (°C)	99.3	99.4
Initial temperature of water (°C)	23.27	23.49
Final temperature of water (°C)	27.14	27.47
$\Delta T$ of water (°C)	3.87	3.98
Energy absorbed by water (J)	$3.24 \times 10^2$	$3.33 \times 10^2$
Specific heat capacity of water (J/g•°C)	4.184	4.184
$\Delta T$ of metal (°C)	-72.2	-71.93
Exp. specific heat capacity of copper (J/g•°C)	0.339	0.343
Actual specific heat capacity of copper (J/g•°C)	0.386 <sup>1</sup>	0.386
Exp. Molar mass of copper (g/mol)	73.7	72.9
Actual molar mass of copper (g/mol)	68.546 <sup>2</sup>	68.546
% Error of specific heat capacity of copper (%)	12.2	11.1
% Error of molar mass of copper (%)	16.0	14.2

### Observations:

- Before the reaction, the water is colourless and the beads of copper are a metallic grey colour
- Once the hot metal is added to the water, the temperature increases quickly and slowly decreases
- After the reaction, the water is still colourless and the beads of copper appear unchanged

<sup>1</sup> <http://hyperphysics.phy-astr.gsu.edu/hbase/tables/sphtt.html>

<sup>2</sup> <http://www.webqc.org/molecular-weight-of-Cu.html>

Table 2: The temperature of water before and after the addition of hot copper

Trial 1		Trial 2	
Time (s)	Temperature (°C)	Time (s)	Temperature (°C)
0.00	23.8	0.00	33.5
15.0	23.6	15.0	24.9
30.0	23.5	30.0	24.1
45.0	23.4	45.0	24.0
60.0	23.4	60.0	23.7
75.0	23.4	75.0	23.6
90.0	23.3	90.0	23.6
105.0	23.2	105.0	23.4
120.0	23.3	120.0	23.5
135.0	23.3	135.0	23.5
150.0	23.2	150.0	23.5
165.0	23.2	165.0	23.4
180.0	23.3	180.0	23.5
Copper Added		Copper Added	
195.0	27.3	195.0	27.5
210.0	27.4	210.0	27.7
225.0	27.3	225.0	27.8
240.0	27.3	240.0	27.6
255.0	27.3	255.0	27.5
270.0	27.2	270.0	27.5
285.0	27.2	285.0	27.5
300.0	27.2	300.0	27.5
315.0	27.2	315.0	27.4
330.0	27.0	330.0	27.3
345.0	27.1	345.0	27.3
360.0	27.1	360.0	27.3
375.0	27.0	375.0	27.3
390.0	27.1	390.0	27.3
405.0	27.1		
420.0	27.0		
435.0	27.0		
450.0	26.9		
465.0	26.9		
480.0	27.0		

Figure 1. Enthalpy of a Metal with copper – Trial 1

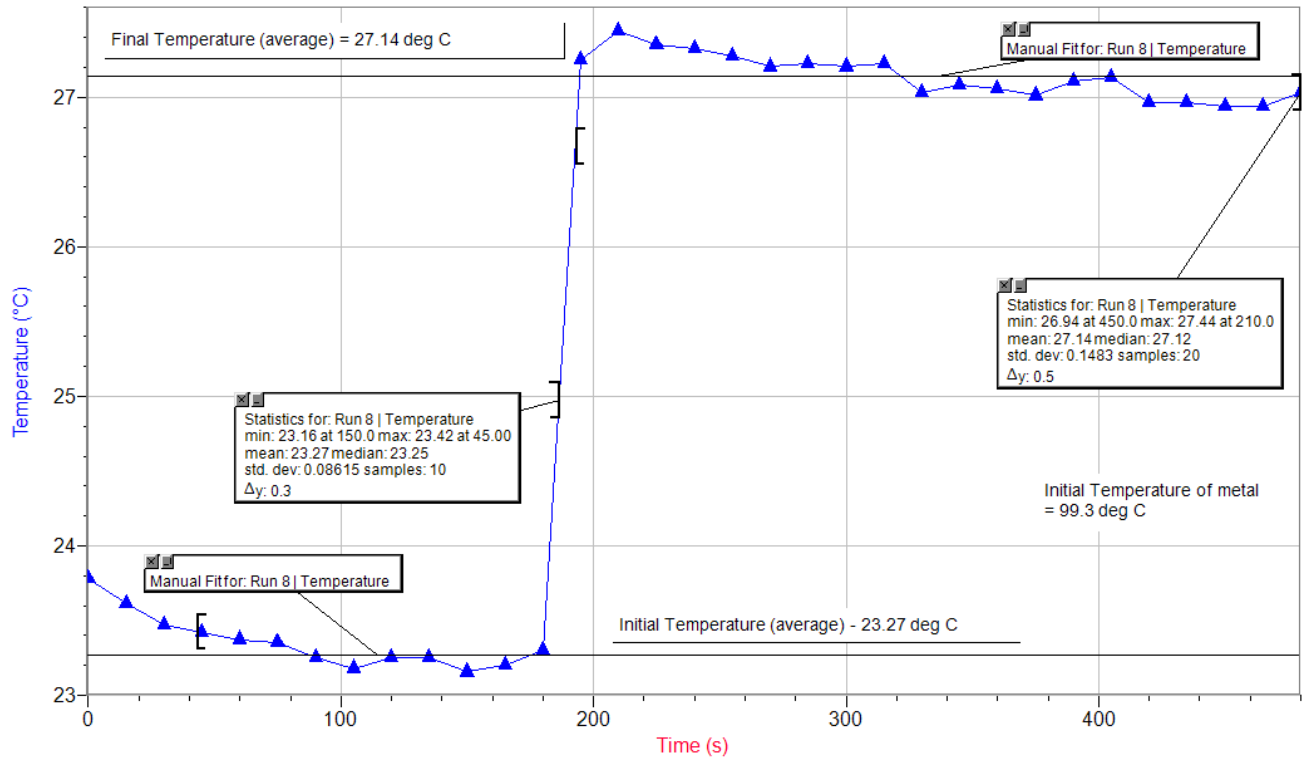


Figure 2. Enthalpy of a metal with copper – Trial 2

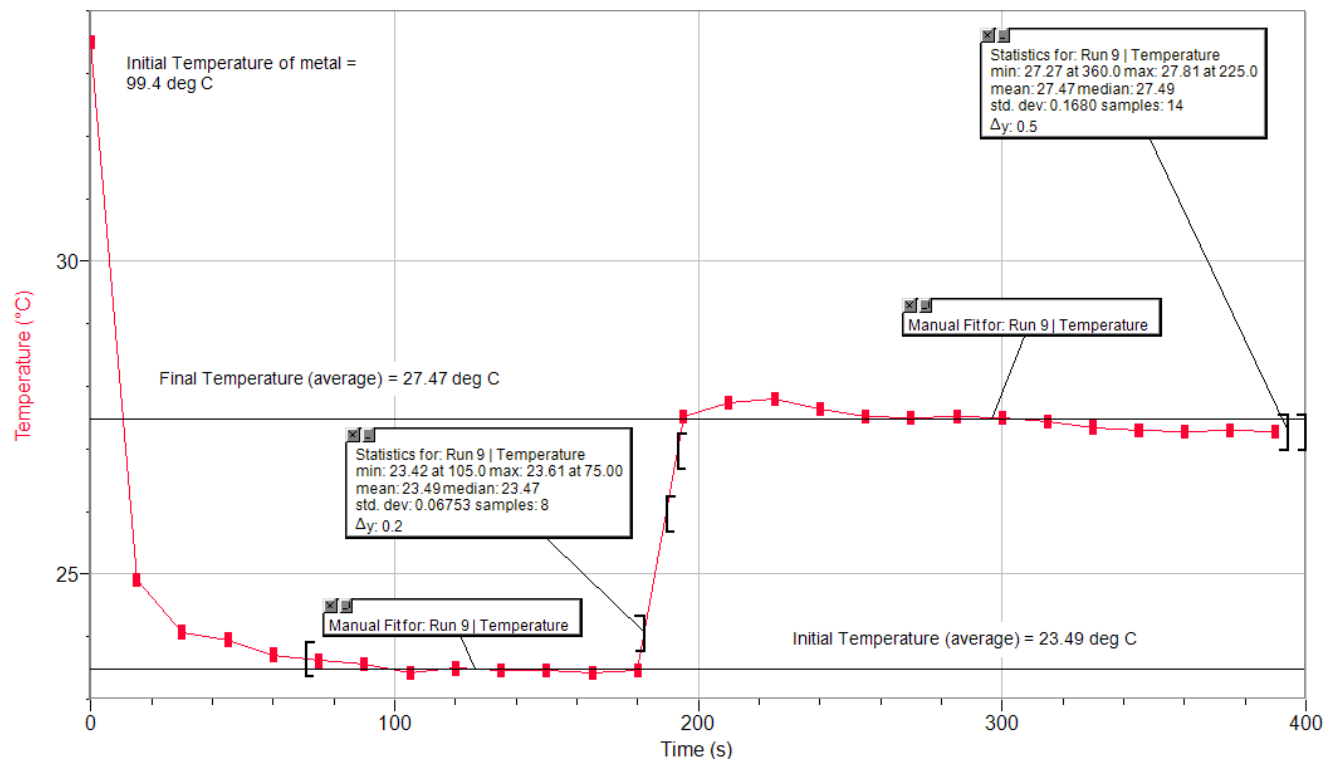


Table 3: Part 2a of the experiment – enthalpy of neutralization using HCl

Data	Trial 1	Trial 2
Base Used	NaOH	NaOH
Concentration of NaOH (mol/L)	1.0	1.0
Volume NaOH (mL)	34.5	34.5
Initial temperature of NaOH (°C)	23.21	22.94
Mass of the empty calorimeter cup (g)	1.8972	1.8972
Acid Used	HCl	HCl
Concentration of HCl (mol/L)	1.1	1.1
Volume HCl (mL)	34.5	34.5
Final temperature of solution (°C)	30.29	29.92
$\Delta T$ of solution (°C)	7.08	6.98
Mass of the calorimeter cup with solution (g)	71.5810	71.3896
Specific heat capacity of solution (J/g•°C)	4.184	4.184
Energy released/absorbed (J)	$-2.06 \times 10^3$	$-2.03 \times 10^3$
Number of moles of OH (mol)	$3.4 \times 10^{-2}$	$3.4 \times 10^{-2}$
Number of moles of H <sub>2</sub> O (mol)	$3.4 \times 10^{-2}$	$3.4 \times 10^{-2}$
Exp. $\Delta H$ of neutralization (J/mol)	$-6.1 \times 10^4$	$-6.0 \times 10^4$
Actual $\Delta H$ of neutralization (kJ/mol)	-57.9 <sup>3</sup>	-57.9
% Error of $\Delta H$ of neutralization (%)	5.4	3.6

Observations:

- Before the reaction, both the acid and the base are colourless
- When the acid is added to the base, the temperature rises quickly then slowly begins to decrease
- After the reaction, the solution appears unchanged

<sup>3</sup>[http://chemwiki.ucdavis.edu/Physical\\_Chemistry/Thermodynamics/State\\_Functions/Enthalpy/Enthalpy\\_Change\\_of\\_Neutralization](http://chemwiki.ucdavis.edu/Physical_Chemistry/Thermodynamics/State_Functions/Enthalpy/Enthalpy_Change_of_Neutralization)

Table 4: The temperature of the solution before and after the addition of HCl

Trial 1		Trial 2	
Time (s)	Temperature (°C)	Time (s)	Temperature (°C)
0.00	23.2	0.00	23.1
15.0	23.3	15.0	23.0
30.0	23.2	30.0	22.9
45.0	23.2	45.0	23.0
60.0	23.1	60.0	23.0
75.0	23.2	75.0	23.0
90.0	23.2	90.0	22.8
105.0	23.2	105.0	22.9
120.0	23.2	120.0	22.9
135.0	23.3	135.0	22.9
150.0	23.2	150.0	23.0
165.0	23.2	165.0	22.9
180.0	23.3	180.0	22.8
HCl added		HCl added	
195.0	30.5	195.0	30.0
210.0	30.4	210.0	30.0
225.0	30.4	225.0	30.2
240.0	30.4	240.0	30.1
255.0	30.4	255.0	30.1
270.0	30.4	270.0	30.0
285.0	30.3	285.0	29.9
300.0	30.2	300.0	30.0
315.0	30.1	315.0	30.0
330.0	30.2	330.0	30.1
345.0	30.2	345.0	29.9
360.0	30.2	360.0	29.8
375.0	30.2	375.0	30.0
390.0	30.2	390.0	29.9
		405.0	29.9
		420.0	30.0
		435.0	29.8
		450.0	29.8
		465.0	29.8
		480.0	29.9
		495.0	29.8
		510.0	29.8
		525.0	29.8
		540.0	29.8

Figure 3. Enthalpy of Neutralization of NaOH and HCl – Trial 1

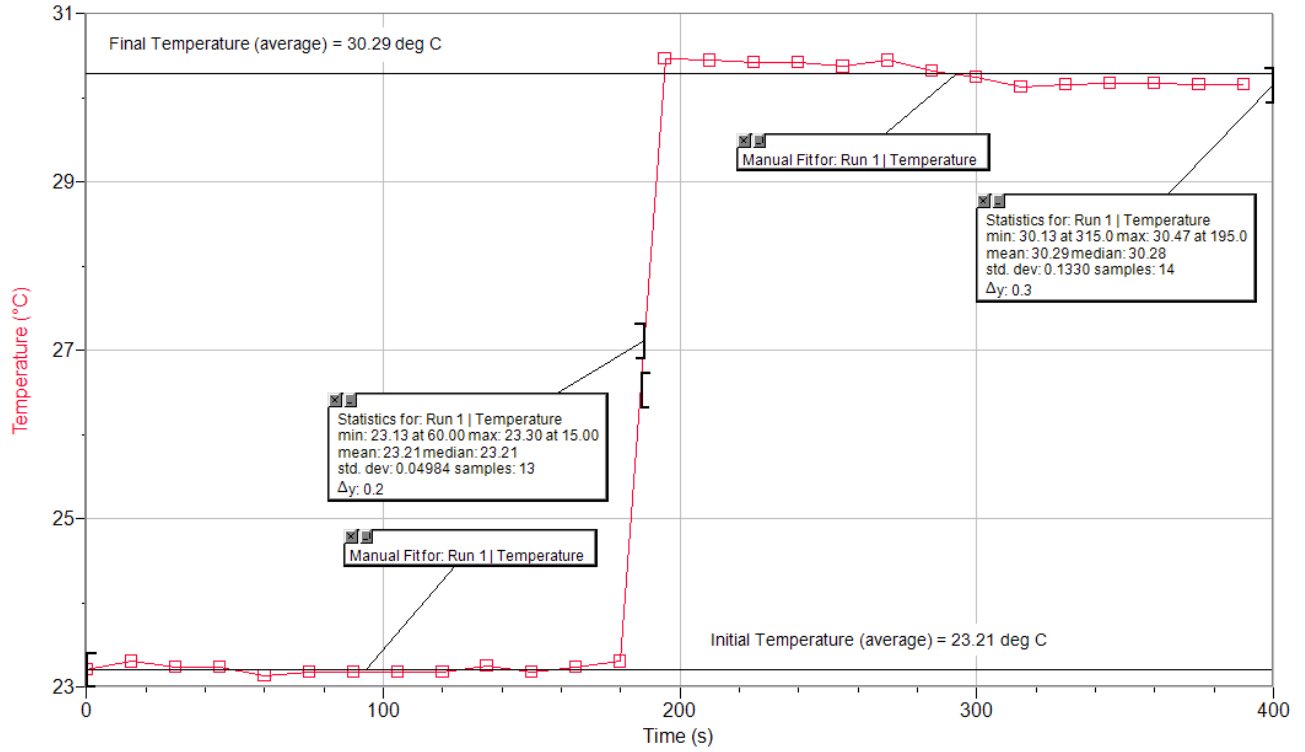


Figure 4. Enthalpy of Neutralization of NaOH and HCl – Trial 2

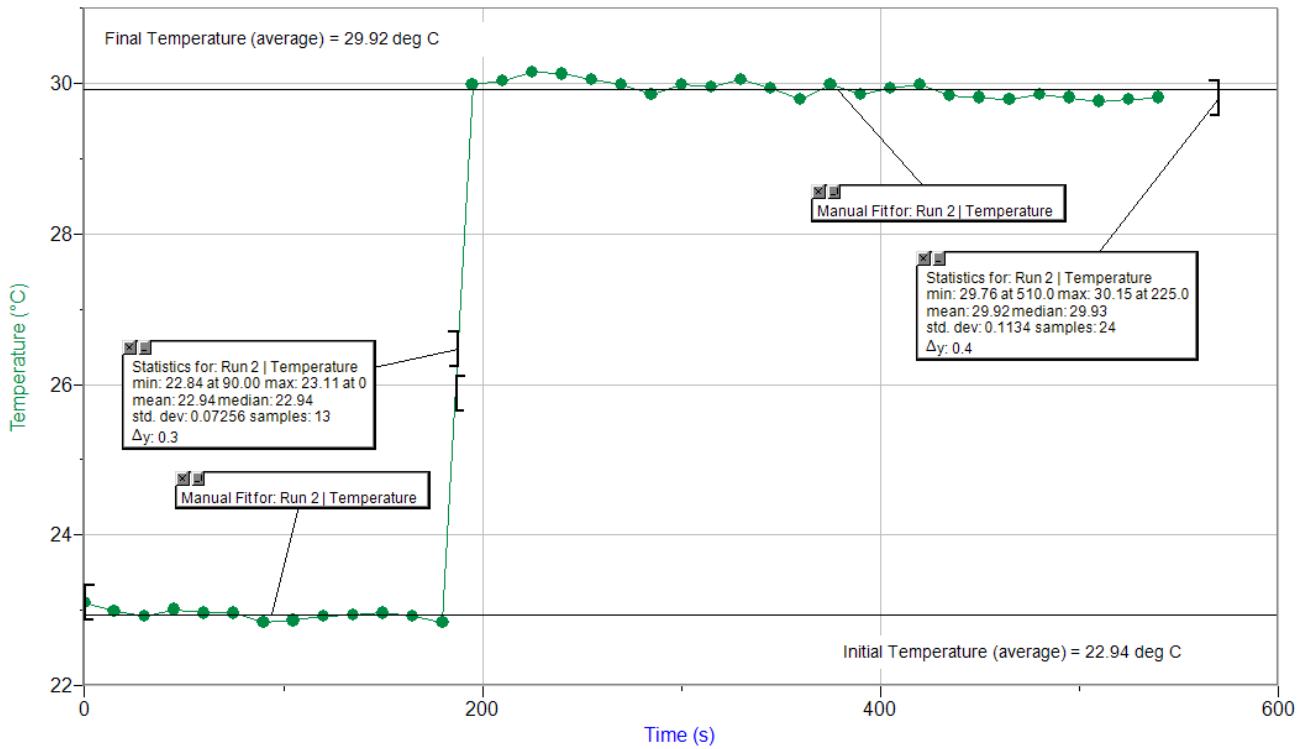


Table 5: Part 2b of the experiment – enthalpy of neutralization using HNO<sub>3</sub>

Data	Trial 1	Trial 2
Base Used	NaOH	NaOH
Concentration of NaOH (mol/L)	1.0	1.0
Volume NaOH (mL)	32.9	46.7
Initial temperature of NaOH (°C)	23.01	23.14
Mass of the empty calorimeter cup (g)	1.8972	1.8972
Acid Used	HNO <sub>3</sub>	HNO <sub>3</sub>
Concentration of HNO <sub>3</sub> (mol/L)	1.1	1.1
Volume HNO <sub>3</sub> (mL)	32.9	32.9
Final temperature of solution (°C)	29.80	29.92
ΔT of solution (°C)	6.79	6.78
Mass of the calorimeter cup with solution (g)	68.7067	93.4
Specific heat capacity of solution (J/g●°C)	4.184	4.184
Energy released (J)	-1.90 x 10 <sup>3</sup>	-2.65 x 10 <sup>3</sup>
Number of moles of OH (mol)	3.3 x 10 <sup>-2</sup>	4.7 x 10 <sup>-2</sup>
Number of moles of H <sub>2</sub> O (mol)	3.3 x 10 <sup>-2</sup>	4.7 x 10 <sup>-2</sup>
Exp. ΔH of neutralization (J/mol)	-5.8 x 10 <sup>4</sup>	-5.6 x 10 <sup>4</sup>
Actual ΔH of neutralization (kJ/mol)	-57.3 <sup>4</sup>	-57.3
% Error of ΔH of neutralization (%)	1.2	2.7

Observations:

- Before the reaction, both the acid and the base are colourless
- When the acid is added to the base, the temperature rises quickly then slowly begins to decrease
- After the reaction, the solution appears unchanged

<sup>4</sup> <http://www.docbrown.info/page07/delta1He.htm>

Table 6: The temperature of the solution before and after the addition of HNO<sub>3</sub>

Trial 1		Trial 2	
Time (s)	Temperature (°C)	Time (s)	Temperature (°C)
0.00	22.8	0.00	22.9
15.0	23.0	15.0	23.1
30.0	23.0	30.0	23.0
45.0	23.0	45.0	23.2
60.0	23.0	60.0	23.2
75.0	23.0	75.0	23.1
90.0	23.0	90.0	23.2
105.0	23.0	105.0	23.2
120.0	23.0	120.0	23.3
135.0	23.0	135.0	23.2
150.0	23.1	150.0	23.2
165.0	23.2	165.0	23.2
180.0	23.0	180.0	23.2
HNO <sub>3</sub> added		HNO <sub>3</sub> added	
195.0	29.8	195.0	29.2
210.0	30.0	210.0	30.1
225.0	29.8	225.0	30.1
240.0	29.9	240.0	30.0
255.0	30.0	255.0	29.9
270.0	29.9	270.0	30.0
285.0	29.8	285.0	30.1
300.0	29.7	300.0	30.0
315.0	29.9	315.0	30.0
330.0	29.8	330.0	29.9
345.0	29.8	345.0	30.0
360.0	29.8	360.0	29.8
375.0	29.9	375.0	29.9
390.0	29.8	390.0	29.9
405.0	29.7	405.0	29.9
420.0	29.6	420.0	29.9
435.0	29.7	435.0	29.9
450.0	29.7	450.0	29.9
465.0	29.7		
480.0	29.7		

Figure 5. Enthalpy of neutralization of NaOH and HNO<sub>3</sub> – Trial 1

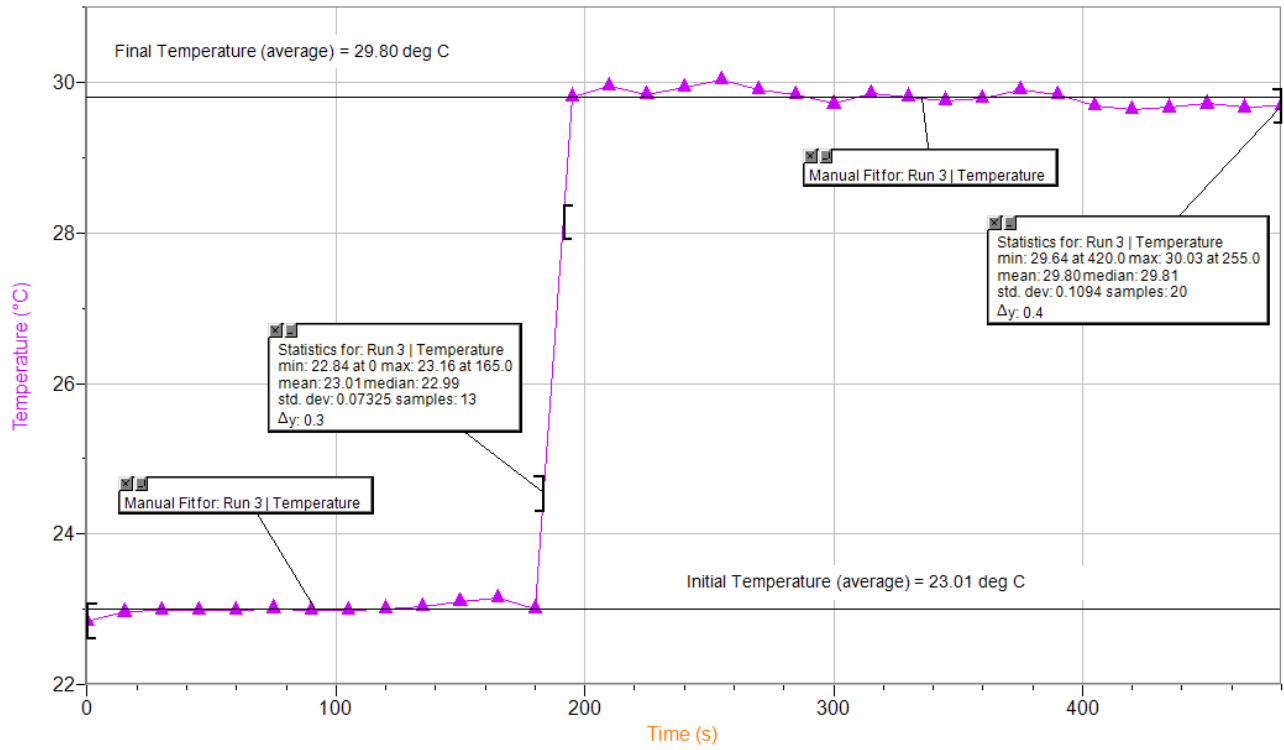


Figure 6. Enthalpy of neutralization of NaOH and HNO<sub>3</sub> – Trial 2

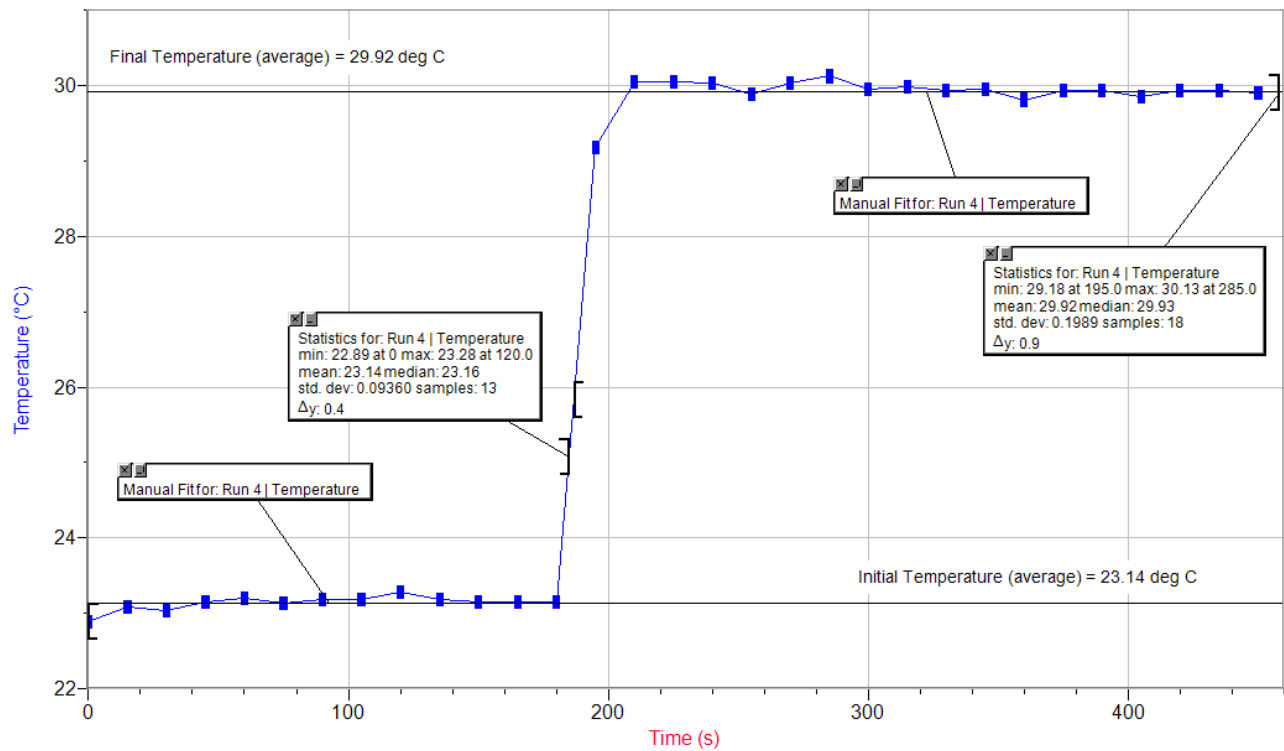


Table 7: Part 3 of the experiment – enthalpy of dissolution of a salt

Data	Trial 1	Trial 2
Unknown Salt	C	C
Mass of salt (g)	1.5009	1.5084
Mass of the empty calorimeter cup (g)	1.8972	1.8972
Volume of water (mL)	20.0	20.0
Mass of the calorimeter cup with water (g)	20.9272	21.2362
Mass of the calorimeter cup with solution (g)	22.4047	22.6133
Mass of solution (g)	20.5075	20.7161
Initial Temperature of Water (°C)	23.09	22.95
Final Temperature of Solution (°C)	18.29	18.02
$\Delta T$ of solution (°C)	-4.80	-4.93
Specific heat capacity of solution (J/g•°C)	4.184	4.184
Energy absorbed (J)	$4.30 \times 10^2$	$4.40 \times 10^2$
Exp. Enthalpy of solution (J/mol)	$2.90 \times 10^4$	$2.95 \times 10^4$
Actual enthalpy of solution (kJ/mol)	34.89 <sup>5</sup>	34.89
% Error of enthalpy of solution (%)	16.9	15.4

Observations:

- The salt is a white powdery substance
- Once the salt is added to the water, the temperature decreases quickly then begins to rise slowly
- After the reaction, the salt has completely dissolved and a colourless solution remains

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<sup>5</sup>[http://sites.chem.colostate.edu/diverdi/all\\_courses/CRC%20reference%20data/enthalpies%20of%20solution%20of%20electrolytes.pdf](http://sites.chem.colostate.edu/diverdi/all_courses/CRC%20reference%20data/enthalpies%20of%20solution%20of%20electrolytes.pdf)

Table 8: The temperature of the solution before and after the addition of Unknown Salt "C"

Trial 1		Trial 2	
Time (s)	Temperature (°C)	Time (s)	Temperature (°C)
0.00	23.0	0.00	22.6
15.0	23.1	15.0	22.9
30.0	23.0	30.0	22.9
45.0	23.0	45.0	23.0
60.0	23.0	60.0	23.0
75.0	23.1	75.0	22.9
90.0	23.1	90.0	23.0
105.0	23.1	105.0	23.0
120.0	23.2	120.0	23.0
135.0	23.2	135.0	23.0
150.0	23.0	150.0	23.1
165.0	23.2	165.0	23.0
180.0	23.2	180.0	23.0
Unknown Salt "C" added		195	22.9
195.0	21.8	Unknown Salt "C" added	
210.0	19.6	210.0	19.0
225.0	18.1	225.0	17.6
240.0	17.8	240.0	17.5
255.0	17.6	255.0	17.5
270.0	17.6	270.0	17.5
285.0	17.8	285.0	17.6
300.0	18.0	300.0	17.6
315.0	18.1	315.0	17.8
330.0	18.0	330.0	17.8
345.0	18.0	345.0	17.6
360.0	18.0	360.0	17.8
375.0	18.1	375.0	17.9
390.0	18.1	390.0	17.8
405.0	18.5	405.0	18.0
420.0	18.2	420.0	18.0
435.0	18.2	435.0	18.1
450.0	18.1	450.0	18.1
465.0	18.3	465.0	18.3
480.0	18.5	480.0	18.2
495.0	18.4	495.0	18.2
510.0	18.4	510.0	18.3
525.0	18.7	525.0	18.4
540.0	18.7	540.0	18.4
555.0	18.7	555.0	18.6
570.0	18.6	570.0	18.5
585.0	18.7	585.0	18.6
600.0	18.7	600.0	18.6

Figure 7. Enthalpy of Dissolution of Unknown Salt C – Trial 1

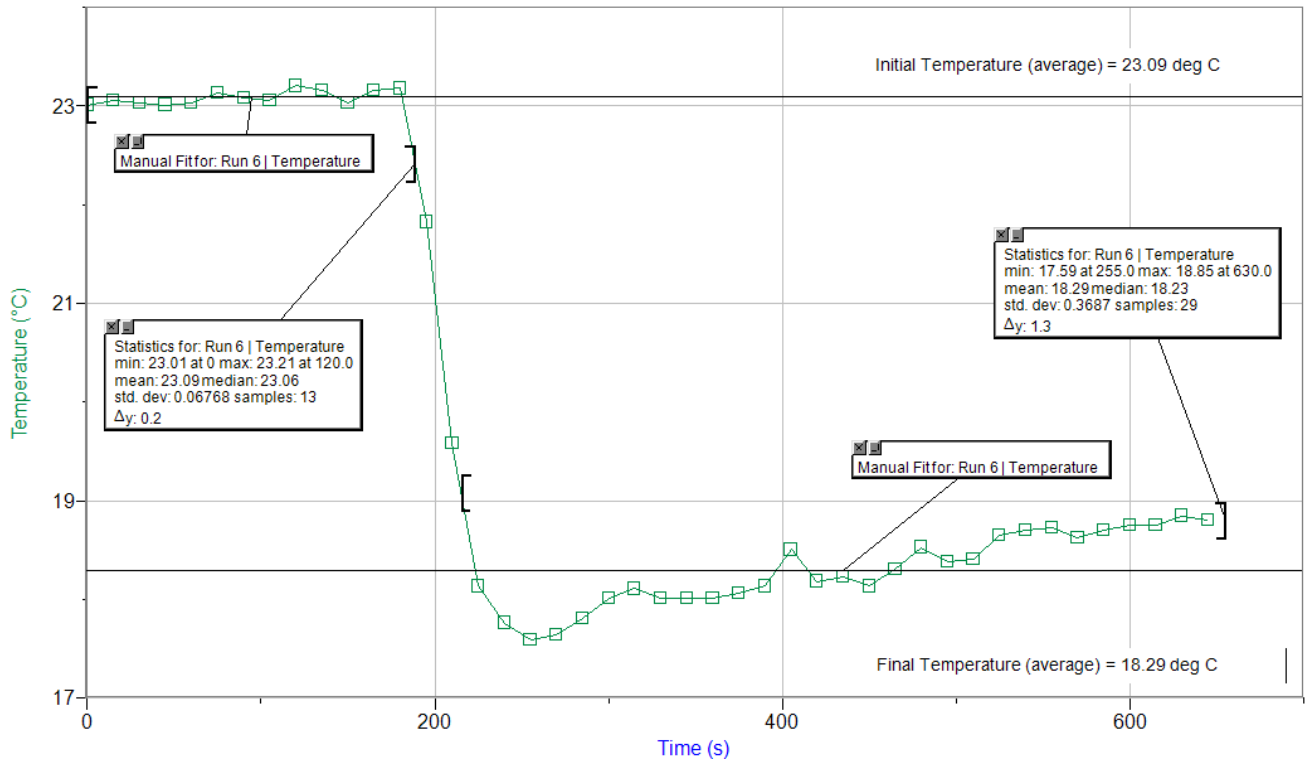
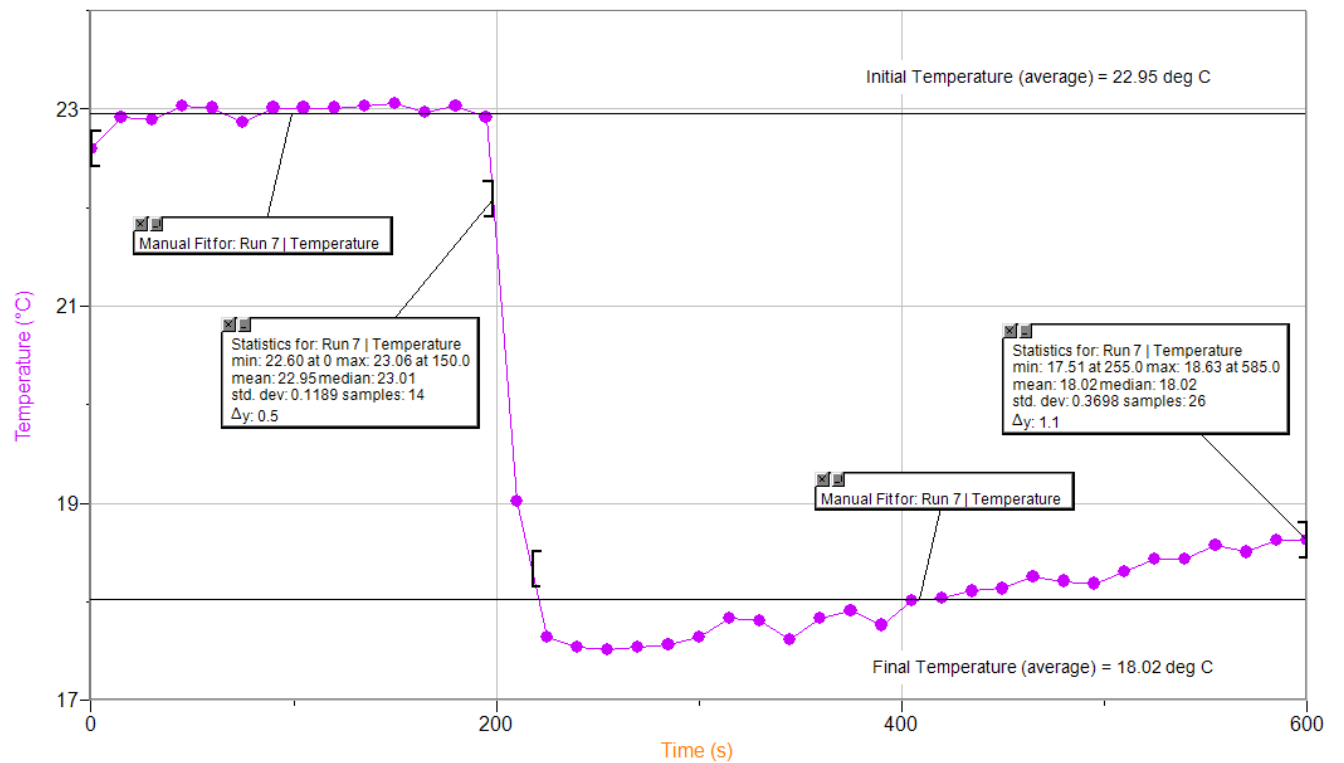


Figure 8. Enthalpy of Dissolution of Unknown Salt C – Trial 2



## Calculations

### Part 1. Enthalpy of a Metal

1.  $\Delta T_{\text{water}} = T_f - T_i$

$$\Delta T_{\text{water}} = 27.14^\circ\text{C} - 23.27^\circ\text{C}$$

$$\Delta T_{\text{water}} = 3.87^\circ\text{C}$$

2.  $q = mc\Delta T$

$$q = (20.0\text{g})(4.184\text{J/g}\cdot^\circ\text{C})(3.87^\circ\text{C})$$

$$q = 3.24 \times 10^2 \text{ J}$$

where  $m = 20.0\text{mL} \times (1.00\text{g}/1.00\text{mL})$

$$m = 20.0 \text{ g}$$

3.  $\Delta T_{\text{metal}} = T_f - T_i$

$$\Delta T_{\text{metal}} = 27.14^\circ\text{C} - 99.3^\circ\text{C}$$

$$\Delta T_{\text{metal}} = -72.2^\circ\text{C}$$

4.  $-q(\text{metal}) = q(\text{water})$

$$-m_{\text{metal}}C_{\text{metal}}\Delta T_{\text{metal}} = m_{\text{water}}C_{\text{water}}\Delta T_{\text{water}}$$

$$C_{\text{metal}} = -m_{\text{water}}C_{\text{water}}\Delta T_{\text{water}} / m_{\text{metal}}\Delta T_{\text{metal}}$$

$$C_{\text{metal}} = - (3.24 \times 10^2 \text{ J}) / (13.2404\text{g})(-72.2^\circ\text{C})$$

$$C_{\text{metal}} = 0.339 \text{ J/g}\cdot^\circ\text{C}$$

5.  $C_{\text{metal}} \times MM_{\text{metal}} = 25 \text{ J/mol}\cdot^\circ\text{C}$

$$MM_{\text{Cu}} = 25 \text{ J/mol}\cdot^\circ\text{C} / C_{\text{Cu}}$$

$$MM_{\text{Cu}} = 25 \text{ J/mol}\cdot^\circ\text{C} / 0.339 \text{ J/g}\cdot^\circ\text{C}$$

$$MM_{\text{Cu}} = 73.7 \text{ g/mol}$$

6. Actual:  $C_{\text{Cu}} = 0.386 \text{ J/g}\cdot^\circ\text{C}$

$$MM_{\text{Cu}} = 63.546 \text{ g/mol}$$

$$\% \text{ error} = (|C_{\text{exp}} - C_{\text{actual}}| / C_{\text{actual}}) \times 100\%$$

$$\% \text{ error} = (|0.339 \text{ J/g}\cdot^\circ\text{C} - 0.386 \text{ J/g}\cdot^\circ\text{C}| / 0.386 \text{ J/g}\cdot^\circ\text{C}) \times 100\%$$

$$\% \text{ error} = 12.2\%$$

$$\% \text{ error} = (|MM_{\text{exp}} - MM_{\text{actual}}| / MM_{\text{actual}}) \times 100\%$$

$$\% \text{ error} = (|73.7 \text{ g/mol} - 63.546 \text{ g/mol}| / 63.546 \text{ g/mol}) \times 100\%$$

$$\% \text{ error} = 16.0\%$$

7. The mass of the metal used made a difference since the mass of the metal has an impact on the amount of heat lost or absorbed by the metal. If the mass of the metal had been lower, there would have been less heat released, and if the mass of the metal had been higher, there would have been more heat released. However, these values do not change the enthalpy of the metal since it takes into account the mass of the metal in its calculation.

### Part 2a. Enthalpy of Neutralization - HCl

$$8. \Delta T_{\text{solution}} = T_{\text{f solution}} - T_{\text{i NaOH}}$$

$$\Delta T_{\text{solution}} = 30.29^{\circ}\text{C} - 23.21^{\circ}\text{C}$$

$$\Delta T_{\text{solution}} = 7.08^{\circ}\text{C}$$

$$9. V_{\text{solution}} = V_{\text{HCl}} + V_{\text{NaOH}}$$

$$V_{\text{solution}} = 34.5 \text{ mL} + 34.5 \text{ mL}$$

$$V_{\text{solution}} = 69.0 \text{ mL}$$

$$10. m_{\text{solution}} = V_{\text{solution}} \times d_{\text{solution}} \qquad m_{\text{solution}} = m_{\text{calorimeter+solution}} - m_{\text{calorimeter}}$$

$$m_{\text{solution}} = 69.0 \text{ mL} \times 1.0 \text{ g/mL} \qquad m_{\text{solution}} = 71.5810 \text{ g} - 1.8972 \text{ g}$$

$$m_{\text{solution}} = 69.0 \text{ g} \qquad m_{\text{solution}} = 69.6838 \text{ g}$$

The two masses are very similar, only differing by a fraction of a gram. The more accurate value is the measured value since that is the mass of the solution remaining after the experiment, it was not derived mathematically. However, it doesn't really matter since the values are so similar.

$$11. q_{\text{N}} = -m_{\text{total}} C_{\text{solution}} \Delta T_{\text{solution}}$$

$$q_{\text{N}} = -(69.6838 \text{ g})(4.184 \text{ J/g}\cdot^{\circ}\text{C})(7.08^{\circ}\text{C})$$

$$q_{\text{N}} = -2.06 \times 10^3 \text{ J}$$

$$12. n_{\text{OH}^-} = V_{\text{NaOH}} \times C_{\text{NaOH}} \times (1 \text{ mol OH}^- / 1 \text{ mol NaOH})$$

$$n_{\text{OH}^-} = (0.0345 \text{ L})(1.0 \text{ mol/L})$$

$$n_{\text{OH}^-} = 3.4 \times 10^{-2} \text{ mol}$$

$$13. \text{ Since the ratio of NaOH to H}_2\text{O is 1:1}$$

$$n_{\text{OH}^-} = 3.4 \times 10^{-2} \text{ mol} = n_{\text{H}_2\text{O}}$$

$$n_{\text{H}_2\text{O}} = 3.4 \times 10^{-2} \text{ mol}$$

$$14. \Delta_N H^\circ = q_N / n_{H_2O}$$

$$\Delta_N H^\circ = (-2.06 \times 10^3 \text{ J}) / (3.4 \times 10^{-2} \text{ mol})$$

$$\Delta_N H^\circ = -6.1 \times 10^4 \text{ J/mol}$$

15. Refer to next page

$$16. \Delta_N H^\circ_{\text{actual}} = -57.9 \text{ kJ/mol}$$

$$\% \text{error} = (|\Delta_N H^\circ_{\text{exp.}} - \Delta_N H^\circ_{\text{actual}}| / \Delta_N H^\circ_{\text{actual}}) \times 100\%$$

$$\% \text{error} = (|-61 \text{ kJ/mol} + 57.9 \text{ kJ/mol}| / -57.9 \text{ kJ/mol}) \times 100\%$$

$$\% \text{error} = 5.4\%$$

17. Refer to next page

### **Part 2b. Enthalpy of Neutralization – HNO<sub>3</sub>**

$$8. \Delta T_{\text{solution}} = T_{f \text{ solution}} - T_{i \text{ NaOH}}$$

$$\Delta T_{\text{solution}} = 29.80^\circ\text{C} - 23.01^\circ\text{C}$$

$$\Delta T_{\text{solution}} = 6.79^\circ\text{C}$$

$$9. V_{\text{solution}} = V_{\text{HCl}} + V_{\text{NaOH}}$$

$$V_{\text{solution}} = 32.9 \text{ mL} + 32.9 \text{ mL}$$

$$V_{\text{solution}} = 65.8 \text{ mL}$$

$$10. m_{\text{solution}} = V_{\text{solution}} \times d_{\text{solution}} \qquad m_{\text{solution}} = m_{\text{calorimeter+solution}} - m_{\text{calorimeter}}$$

$$m_{\text{solution}} = 65.8 \text{ mL} \times 1.0 \text{ g/mL} \qquad m_{\text{solution}} = 68.7067 \text{ g} - 1.8972 \text{ g}$$

$$m_{\text{solution}} = 65.8 \text{ g} \qquad m_{\text{solution}} = 66.8095 \text{ g}$$

The two masses are very similar, only differing by only about one gram. The more accurate value is the measured value since that is the mass of the solution remaining after the experiment, it was not derived mathematically. However, it doesn't really matter since the values are so similar.

$$11. q_N = -m_{\text{total}} C_{\text{solution}} \Delta T_{\text{solution}}$$

$$q_N = -(66.8095 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C})(6.79^\circ\text{C})$$

$$q_N = -1.90 \times 10^3 \text{ J}$$

12.  $n_{\text{OH}^-} = V_{\text{NaOH}} \times C_{\text{NaOH}} \times (1 \text{ mol OH}^- / 1 \text{ mol NaOH})$

$$n_{\text{OH}^-} = (0.0329 \text{ L})(1.0 \text{ mol/L})$$

$$n_{\text{OH}^-} = 3.3 \times 10^{-2} \text{ mol}$$

13. Since the ratio of NaOH to H<sub>2</sub>O is 1:1

$$n_{\text{OH}^-} = 3.3 \times 10^{-2} \text{ mol} = n_{\text{H}_2\text{O}}$$

$$n_{\text{H}_2\text{O}} = 3.3 \times 10^{-2} \text{ mol}$$

14.  $\Delta_N H^\circ = q_N / n_{\text{H}_2\text{O}}$

$$\Delta_N H^\circ = (-1.90 \times 10^3 \text{ J}) / (3.3 \times 10^{-2} \text{ mol})$$

$$\Delta_N H^\circ = -5.8 \times 10^4 \text{ J/mol}$$

15. The experimental heat of neutralization per mole of water with the strong acid HCl is  $-6.1 \times 10^4$  J/mol, and the experimental heat of neutralization per mole of water with the strong acid HNO<sub>3</sub> is  $-5.8 \times 10^4$  J. The two values are similar to one another, varying only by about 3 kJ.

16.  $\Delta_N H^\circ_{\text{actual}} = -57.3 \text{ kJ/mol}$

$$\% \text{error} = (|\Delta_N H^\circ_{\text{exp.}} - \Delta_N H^\circ_{\text{actual}}| / \Delta_N H^\circ_{\text{actual}}) \times 100\%$$

$$\% \text{error} = (|-58 \text{ kJ/mol} + 57.3 \text{ kJ/mol}| / -57.3 \text{ kJ/mol}) \times 100\%$$

$$\% \text{error} = 1.2\%$$

17. For HCl, the experimental enthalpies of neutralizations were  $-6.1 \times 10^4$  J/mol, and  $-6.0 \times 10^4$  J/mol respectively. The values are very precise as they only vary by approximately 1 kJ. For HNO<sub>3</sub>, the experimental enthalpies of neutralizations were  $-5.8 \times 10^4$  J/mol and  $-5.6 \times 10^4$  J/mol respectively. The values are also quite precise as they only vary by about 2 kJ. The volume for the HCl was the same for both trials, but it varied with HNO<sub>3</sub>. The greater the volume of the base, the greater the mass of the solution, and the greater the number of moles of OH<sup>-</sup> and H<sub>2</sub>O. However, a variation in volume does not affect the enthalpy of neutralization, nor should it because the enthalpy is measured in J/mol so a greater volume would yield a greater number of joules but also a greater number of moles. The values of enthalpy of neutralization for the two different acids were quite similar as well ranging only by 2-5 kJ. I did expect the results that I obtained (in reference to the similar results for the two different acids) because a neutralization reaction is basically producing water from H<sup>+</sup> and OH<sup>-</sup>, regardless of the strong acid used, therefore the enthalpies of neutralization should be similar.

### Part 3. Enthalpy of Dissolution of a Salt

$$\begin{aligned}18. \Delta T_{\text{solution}} &= T_{\text{f solution}} - T_{\text{i water}} \\ \Delta T_{\text{solution}} &= 18.29^{\circ}\text{C} - 23.09^{\circ}\text{C} \\ \Delta T_{\text{solution}} &= -4.80^{\circ}\text{C}\end{aligned}$$

$$\begin{aligned}19. q_{\text{solution}} &= -q_{\text{water}} - q_{\text{salt}} \\ q_{\text{solution}} &= -(m_{\text{water}}c_{\text{water}}\Delta T_{\text{water}}) - (m_{\text{salt}}c_{\text{salt}}\Delta T_{\text{salt}}) \\ q_{\text{solution}} &= -(20.0 \text{ g})(4.184 \text{ J/g}\cdot^{\circ}\text{C})(-4.80^{\circ}\text{C}) - (1.5009 \text{ g})(3.877 \text{ J/g}\cdot^{\circ}\text{C})(-4.80^{\circ}\text{C}) \\ q_{\text{solution}} &= 4.30 \times 10^2 \text{ J}\end{aligned}$$

$$\begin{aligned}20. \Delta_{\text{s}}H^{\circ} &= q_{\text{solution}} / n_{\text{salt}} & \text{where } n_{\text{salt}} &= m_{\text{salt}}/MM_{\text{salt}} \\ \Delta_{\text{s}}H^{\circ} &= (4.30 \times 10^2 \text{ J}) / 0.014844 \text{ mol} & n_{\text{salt}} &= (1.5009 \text{ g}) / (101.11 \text{ g/mol}) \\ \Delta_{\text{s}}H^{\circ} &= 2.90 \times 10^4 \text{ J/mol} & n_{\text{salt}} &= 0.014844 \text{ mol}\end{aligned}$$

21. Potassium nitrate has a molar mass of  $101.1 \text{ g/mol}^6$ , which suggests that Unknown Salt C is potassium nitrate.

$$\Delta_{\text{s}}H^{\circ}_{\text{actual}} = 34.89 \text{ kJ/mol}$$

$$\% \text{error} = (|\Delta_{\text{s}}H^{\circ}_{\text{exp.}} - \Delta_{\text{s}}H^{\circ}_{\text{actual}}| / \Delta_{\text{s}}H^{\circ}_{\text{actual}}) \times 100\%$$

$$\% \text{error} = (|29.0 \text{ kJ/mol} - 34.89 \text{ kJ/mol}| / 34.89 \text{ kJ/mol}) \times 100\%$$

$$\% \text{error} = 16.9\%$$

22. The mass of salt did affect the results. A smaller mass of salt would have caused the energy absorbed by the solution smaller. However, this would not impact the enthalpy of dissolution since the number of moles of salt is accounted for as the heat of solution is divided by the number of moles of salt.

### Discussion

In the first part of the experiment, the results were precise but not very accurate. The experimental values for the specific heat capacity were found to be  $0.339 \text{ J/g}\cdot^{\circ}\text{C}$  and  $0.343 \text{ J/g}\cdot^{\circ}\text{C}$ . These values were reasonable, however they were a bit low, with percent errors of 12.2% and 11.1% respectively. The experimental molar masses were found to be  $73.7 \text{ g/mol}$  and  $72.9 \text{ g/mol}$  with percent errors of 16.0% and 14.2%. In the determination of the initial temperature, the first three points on the graph were omitted from the analysis. This is because the temperature of the boiling water had been measured

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<sup>6</sup> [http://pubchem.ncbi.nlm.nih.gov/compound/potassium\\_nitrate](http://pubchem.ncbi.nlm.nih.gov/compound/potassium_nitrate)

right before, and the probe had not completely cooled down. By including these first values, the accuracy of the initial temperature would have been lower.

In the second part of the experiment, the results were both very precise and accurate. The experimental enthalpies of neutralization using HCl were found to be  $-6.1 \times 10^4$  J/mol and  $-6.0 \times 10^4$  J/mol. The experimental enthalpies of neutralization using HNO<sub>3</sub> were found to be  $-5.8 \times 10^4$  J/mol and  $-5.6 \times 10^4$  J/mol. These values are quite similar, ranging from only 2-5 kJ. The percent errors of the two acids were 5.4% and 3.6% for HCl and 1.2% and 2.7% for HNO<sub>3</sub>. The literature values for these enthalpies range from -57 kJ/mol to -58 kJ/mol. The values for enthalpies of neutralization using strong acids fall within the same range since a neutralization reaction involves the production of water from H<sup>+</sup> and OH<sup>-</sup>, which is why these values are so similar.

In the third part of the experiment, the results were very precise, but not very accurate. The experimental enthalpies of the salt were found to be  $2.90 \times 10^4$  J/mol and  $2.95 \times 10^4$  J/mol with percent errors of 16.9% and 15.4% respectively. The identity of the salt was found to be KNO<sub>3</sub> (refer to question 21). The value of the enthalpy of solution (34.89 kJ/mol) was about 5 kJ/mol greater than the experimental value, thus indicating that there was some error in the experiment.

None of the values determined in this experiment were completely accurate, and so there were a few sources of error in the lab. Firstly, to speed up the reaction, the calorimeter was being swirled. This means that some of the recorded values of temperature were slightly too high or too low, thus impacting the accuracy of the results. Additionally, the thermometer probe reached the bottom of the calorimeter cup, which means that when the hot metal was added, the temperature recorded would have been greater than that of the solution. This means that when taking the average of the points, the final temperature may have been too large. With a larger final temperature, the temperature change may also have been too large, thus resulting in more inaccurate results. Another inherent source of error is the heat leaking in or out of the calorimeter. The calorimeter has some holes in the lid and so it is not a perfectly isolated system. Furthermore, when mixing the two substances together, the lid was removed thus allowing heat to escape or enter the calorimeter. This means that the temperature of the substance would slowly decrease for the first two parts, and slowly increase for the third part, rather than stabilizing at the most accurate value. Therefore, when calculating the final temperature using the average, it may be too large or too small, causing the results to be less accurate.

## **Conclusion**

The experimental specific heat capacities of copper were found to be  $0.339$  J/g•°C and  $0.343$  J/g•°C with percent errors of 12.2% and 11.1%, while the experimental molar masses were found to be  $73.7$  g/mol and  $72.9$  g/mol with percent errors of 16.0% and 14.2%. The experimental enthalpies of neutralization with HCl were found to be  $-6.1 \times 10^4$  J/mol and  $-6.0 \times 10^4$  J/mol with percent errors of 5.4% and 3.6%. The experimental enthalpies of neutralization with HNO<sub>3</sub> are found to be  $-5.8 \times 10^4$  J/mol and  $-5.6$  J/mol with percent errors of 1.2% and 2.7%. Finally, the experimental values for the enthalpy of solution of the unknown salt – found to be KNO<sub>3</sub> – is found to be  $2.90 \times 10^4$  J/mol and  $2.95 \times 10^4$  J/mol with percent errors of 16.9% and 15.4%.

B1

C NaOH  $\rightarrow$  1.0 mol/L  
V NaOH  $\rightarrow$  34.5 mL

HCl { C Acid  $\rightarrow$  1.1 mol/L  
V Acid  $\rightarrow$  34.5 mL

m cup + sol'n = 71.5810g  
m cup = 1.8972g

B3

C NaOH  $\rightarrow$  1.0 mol/L  
V NaOH  $\rightarrow$  32.9 mL

HNO<sub>3</sub> { C Acid  $\rightarrow$  1.1 mol/L  
V Acid  $\rightarrow$  32.9 mL

m cup + sol'n = ~~70.0000~~ 68.7067g

unknown salt "C"

C1 m (should be close to 1.5g) = ~~1.5096g~~ 1.5009g  
V water = 20.0 mL  
m water + cup = ~~21.0393g~~ 20.9272g  
m cup + sol'n = 22.4047g

A) Copper

8-15g) m = 13.2404g  
V water = 20.0 mL  
m cup + water = 21.1836g  
T metal ini = 99.3°C

B2

C NaOH  $\rightarrow$  1.0 mol/L  
V NaOH  $\rightarrow$  34.5

HCl { C Acid  $\rightarrow$  1.1 mol/L  
V Acid  $\rightarrow$  34.5 mL

~~m cup + sol'n =~~  
m cup + sol'n = 71.3896g

B4

C NaOH  $\rightarrow$  1.0 mol/L  
V NaOH  $\rightarrow$  46.7 mL

HNO<sub>3</sub> { C Acid  $\rightarrow$  1.1 mol/L  
V Acid  $\rightarrow$  46.7 mL

~~m cup + sol'n =~~

C2 m salt = 1.5084g  
V water = 20.0 mL  
m water + cup = 21.2362g  
m cup + sol'n = 22.6133g

A) Copper

2) m = 13.49  
V water = 20.0 mL  
m cup + water = 21.1070g  
T metal ini = 99.4°C