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PLEASE NOTE: If ANY of the above information is UNCLEAR or not provided, your grade will NOT be recorded!!

Lab Day (circle): Tues aft Tues night Wed Thurs aft Thurs night Fri

Lab Week (circle): 1 2

Laboratory Report Cover Page

Experiment 2.

Enthalpy of Various Reactions

Checklist:

- **Raw Data Sheet written in pen, signed by TA and attached**
- **8 curves [2 for metal; 2 for acid 1; 2 for acid 2; 2 for salt] attached**
- **Completed formal report typed and attached**

Student's Initials: ZH

Introduction:

In this experiment, several experiments will be held in order to calculate the enthalpies of various reactions, such as the enthalpy of a metal, enthalpy of neutralization and enthalpy of dissolution of a salt.

A calorimeter is used to measure the thermal energy gained or lost by the chemical reaction. In this experiment, a simple calorimeter, also known as the “coffee cup calorimeter”, is being used. A simple calorimeter is constructed using two nested Styrofoam cups with lids and a thermometer. Although not strictly true, an assumption has been made that in this calorimeter, no heat is lost to the surroundings or the calorimeter as there is no such thing as a perfectly isolated system.

To measure thermal energy, we need to measure the specific heat capacity, which is the amount of thermal energy needed to heat one gram of a substance by one degree. We can determine the energy required to change the temperature of a substance can be determined using the following equation:

$$q=mc\Delta T$$

The c represents the specific heat capacity of a substance. The m is the mass and ΔT is the change in temperature. When the hot metal is added to the water, the temperature of the metal decreases and the temperature of water rises. Therefore, the heat lost by the metal is equal to the heat gained by the water:

$$-q(metal)=q(water) \quad \text{Thus;}$$

$$- mc\Delta T(metal)= mc\Delta T(water)$$

It is possible to calculate the molar mass of a metal once its specific heat capacity is known using the equation:

$$c(metal) \times MM(metal) = 25 \text{ J/mol.C}$$

As for the enthalpy of neutralization, the same principle is used ($q= mc\Delta T$), where q in this case is the total heat energy of the solution, the m is the total mass of the solution, and c is the same as water (4.18 J/mol C). Once the amount of heat released is calculated, the enthalpy of the reaction per mole is found using:

$$\Delta H= q/n$$

Moreover, the dissolution of a salt can be exothermic or endothermic depending on two factors; the lattice energy and the hydration energy. Both these quantities are difficult to measure in a lab. However, their sum, the enthalpy of solution, can be measured without much difficulty. The enthalpy of solution can be expressed as:

$$\Delta H = (\text{-heat loss of } H_2O) + (\text{- heat loss of salt})$$

However in this lab, we will calculate the enthalpy of solution using the following the equation below.

$$\Delta H = q/n = - (m)(c)(\Delta T)/n$$

Part 1 Raw Data

Metal (Cu)

	Trial 1	**Trial 2**
Mass of metal (g)	11.7108	12.0252
Mass of foam cup (g)	2.1231	
Volume of water (ml)	20	
Mass of water + foam cup (g)	20.766	20.649

Observations:

Graph spiked after adding hot Copper

Water is colourless and remains colourless after mixing with metal

Copper residue left inside the calorimeter

Part 2 Raw Data

A) HCl + NaOH

	Trial 1	**Trial 2**
Volume of NaOH (ml)	40	40
Concentration of NaOH (M)	1	1
Volume of HCl (ml)	40	40
Concentration of HCl (M)	1.1	1.1
Mass of foam cup with solution after reaction (g)	66.7135	66.8019

B) HNO_3 + NaOH

	Trial 1	**Trial 2**
Volume of NaOH (ml)	40	40
Concentration of NaOH (M)	1	1
Volume of HNO_3 (ml)	40	40
Concentration of HNO_3 (M)	1.1	1.1
Mass of foam cup with solution after reaction (g)	75.9210	76.1346

Observations:
Graph spiked when acid was added to the base
No fumes from acid
Acid and base are colourless

Part 3 Raw Data

Salt A

	Trial 1	**Trial 2**
Mass of salt (g)	1.4778	1.4959
Volume of water (ml)	20	
Mass of solution + foam cup after mixing (g)	22.4344	21.4038

Observations:

Salt is white and odorless

Drop in graph when salt is added to the water.

Calculations

Part 1.

1) ΔT of water = T2-T1 = 26.7 – 23.1
   = 3.6 C

2) $Q=mc\Delta T$;m= mass of water & foam cup – (mass of Copper+foam cup)
   = 20 x 4.184 x 3.6
   = 301.24 J

3) ΔT_{metal} = Tf – Ti
   = 26.7 – 100
   = -73.3 C

4) $C= q/m\Delta T$
   = -301.12/11.71 x (-73.3)
   = 0.351 J/mol C

5) c x Mr = 25
   Mr= 25/0.351
   = 71.22 g/mol

6) %error (Mr)= (|accepted value – calculated value|/accepted value)x100%
   = (|63.54 – 71.22|/63) x100%
   = 12.1%

   % for C error = ((theoretical value – measured value)/(theoretical value))*100
   = ((0.386-0.351)/(0.386)) x100%
   =9.067%

Part 2

1) $\Delta T = Tf – Ti$
   = 30.46 – 23.40
   = 7.06 C

2) Volume of Acid + Volume of Base=

   40 + 40 = 80 ml

3) Mass = pV
   = 80 x 1
   = 80g

4) $q=mc\Delta T$
   =80x4.184x7.06
   =2363.1232 J

5) HCl(aq) + NaOH(aq) → NaCl(aq) + H2O
   H+(aq) + Cl–(aq) + Na+(aq) + OH–(aq) → Na+(aq) + Cl–(aq) + H2O
   H+(aq) + OH–(aq) → H2O

   Mol of NaOH= 1 mol/ 1000ml = x/40ml
   x= 0.04 mol
   Mol of HCL= 1.1/1000ml= x/40ml
   x= 0.044 mol

   Because it is 1:1:1 ratio with HCl NaOH and OH-, the moles of the limiting reagent OH- = 0.04mol

6) Moles of water = moles OH- (1:1 ratio)
   = 0.04 mol

7) $\Delta H = -q/n$
   = -2363.1232/0.04
   = -59078.08 J/mol
   = -59.07808 KJ/mol

8) %error HCl = |accepted value – calculated value|/ accepted)x100%
   = |-57.7-(-59.08)|/57.7 x100%
   = 2.39 %
   %error HNO_3= |-55.8(-54.308)|55.8 x100%

=5.87 %

Part 3

1) $\Delta T = T2 - T1$
   $=19.84 - 23.80$
   $= -3.96$ C

2) $q=mc\Delta T$
   $=1.4778 \times 3.815 \times -3.96$
   $= -22.32$ J

3) $\Delta H = q/n$

n=m/Mr
= (1.4778)/74.55
= 0.0198 moles of Salt A

Therefore
$\Delta H= (-20.3113)(-3.96) / 0.0198$
$= 15.49$ KJ/mol

4) The salt is KCl (A salt with a Mr of 74.55)

Literature value of ΔH dissolution = 17.57 KJ/mol

%error = (17.57-15.49)/17.57 x100%
= 11.8%

Discussion:

There are a lot of errors that affected the results in this experiment. The most contributing factor is the fact that there is no such thing as a perfectly isolated system; heat loss to the colorimeter and surroundings. This error can been seen through the graphs. In all exothermic reactions' graphs, the temperature was slowly dropping. This will keep on happening until the temperature of the system is at equilibrium with room temperature. As for endothermic reactions, the temperature after dropping will keep on rising until it reaches room temperature. The points on the metal reaction with water graphs (run1 and run2) are fluctuated after mixing because the heat transfer from the metal to the water is not distributed equally among the water. This can alter the final temperature of the water. Another error is the assumption that the solution after mixing the acid and base has a density of 1. This affects the final mass of the solution. The best way to reduce all errors is to do multiple trials of each experiment and take the average of them.

For part 1, the mass of the metal does not affect the final result, as the specific heat capacity of a substance is constant always. Moreover, for part 2, the HCL slightly produces more heat when it reacts with NaOH. The volume of the base would only affect the reaction if its volume was less than the HCl. As for part 3, the mass of the salt wouldn't have affected the result as enthalpy is constant for every substance.

Conclusion:

The percentage error of the enthalpy of Copper in water using the molar mass was 12.1%, and 9.067% when using the specific heat capacity.

The Enthalpy of HCl when mixed with NaOH is -59.07808 KJ/mol, and it had a 2.39 % percentage error away from the literature value.

Salt A in this case was KCl (using the molar mass). Its enthalpy of dissolution is 15.49 KJ/mol and it was 11.80% away from the literature value.