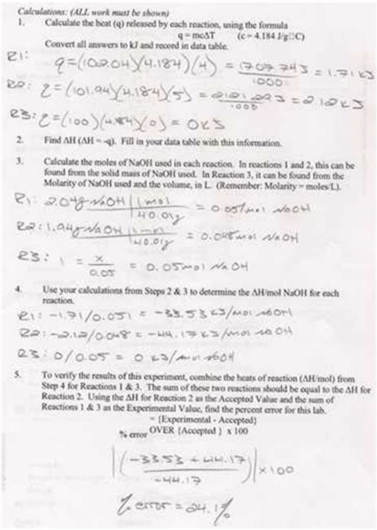

Calculations: (ALL work must be shown)

1. Calculate the heat (q) released by each reaction, using the formula

$q = mc\Delta T$ ($c = 4.184$ J/g□C)

Convert all answers to kJ and record in data table.

R1: $q = (102.04)(4.184)(4) = \frac{1707.74\,J}{1000} = 1.71\,kJ$

R2: $q = (101.94)(4.184)(5) = \frac{2121.22\,J}{1000} = 2.12\,kJ$

R3: $q = (100)(4.184)(0) = 0\,kJ$

2. Find ΔH (ΔH = -q). Fill in your data table with this information.

3. Calculate the moles of NaOH used in each reaction. In reactions 1 and 2, this can be found from the solid mass of NaOH used. In Reaction 3, it can be found from the Molarity of NaOH used and the volume, in L. (Remember: Molarity = moles/L).

R1: 2.04 g NaOH | 1 mol / 40.01 g = 0.051 mol NaOH

R2: 1.94 g NaOH | 1 mol / 40.01 g = 0.048 mol NaOH

R3: $1 = \frac{x}{0.05}$ = 0.05 mol NaOH

4. Use your calculations from Steps 2 & 3 to determine the ΔH/mol NaOH for each reaction.

R1: −1.71/0.051 = −33.53 kJ/mol NaOH

R2: −2.12/0.048 = −44.17 kJ/mol NaOH

R3: 0/0.05 = 0 kJ/mol NaOH

5. To verify the results of this experiment, combine the heats of reaction (ΔH/mol) from Step 4 for Reactions 1 & 3. The sum of these two reactions should be equal to the ΔH for Reaction 2. Using the ΔH for Reaction 2 as the Accepted Value and the sum of Reactions 1 & 3 as the Experimental Value, find the percent error for this lab.

% error = (Experimental - Accepted) OVER (Accepted) x 100

$\left|\left(\frac{-33.53 + 44.17}{-44.17}\right)\right| \times 100$

% error = 24.1%