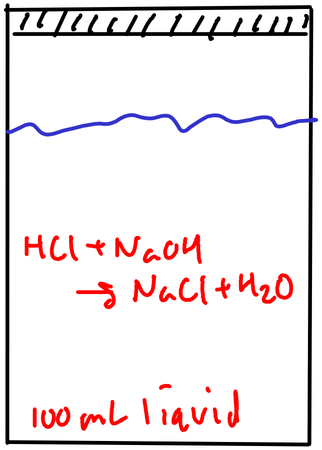
Calorimetry experiment



50 mL 2 M HCl

50 mL 2 M NaOH

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100 mL total volume

We will add 50 mL of 2.0 M NaOH solution tp 50 mL of 2.0 M HCl solution in this experiment!

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

We wish to find the enthalpy change per mole of hydrochloric acid for the reaction above.

We define the system as the reactants and products, and the surroundings as the water in the original two solutions and the cups containing the reaction.

$$\overset{\text{SYSTEM}}{Q_{rxn}} + \overbrace{Q_{water} + Q_{cup}}^{\text{SURROUNDINGS}} = 0$$

First law of thermodynamics

$$Q_{cup} = \Delta T \times 10.\ J/{}^{\circ}C$$

approximate heat capacity of syrofoam cups

$$Q_{water} = 4,184 \frac{J}{g^{\circ}C} \times (100 g) \times \Delta T$$

We will assume the liquid volume is mostly waten and has a specific heat equal to water.

We mix the solutions together and monitor the temperature of the liquid in the cup. Heat loss actually begins the moment the temperature rises higher than room temperature, so we monitor the cup over a period of time to see what the rate of heat loss is. We can use this information to correct our data for heat loss if necessary.

EXPERIMENTAL DATA

before mixing

T_i: 23.8 °C

T_{30} 28.0 °C

T_{60} 28.0 °C

T_{90} 28.0 °C

T_{120} 27.9 °C

T_{150} 27.9 °C

T_{180} 27.9 °C

T_{210} 27.9 °C

T_{240} 27.9 °C

T_f 28.0 °C

The classroom was quite warm, and we saw a slow rate of heat transfer from the cup to the room!

$\Delta T = T_f - T_i =$ 4.2 °C

$Q_{cup} = 10.\ J/°C \times$ 4.2 °C = 42 J

$Q_{water} = 4.184\ J/g°C \times 100.g \times$ 4.2 °C =

= 1757.28 J

$Q_{rxn} = -(\overset{Q_w}{1757.28}\ J + \overset{Q_{cup}}{42}\ J)$

= −1799.28 J

Now, we need to find the enthalpy change per mole of hydrochloric acid

Enthalpy change is equal to heat change at constant pressure, and this was a constant pressure experiment.

$$\Delta H = \frac{Q_{rxn}}{\text{mol HCl}}$$

50.0 mL of 2.0 M HCl was used

$$0.0500\,\text{L} \times 2\,\frac{\text{mol}}{\text{L}} = 0.100\ \text{mol HCl}$$

$$\Delta H = \frac{-1799.28\,\text{J}}{0.100\,\text{mol}} = \frac{-17992.8}{}\ \text{J/mol} = \frac{-18.}{}\ \frac{\text{kJ}}{\text{mol}}$$

So the enthalpy change per mole HCl was determined to be -18 kJ/mol.

This is a good deal lower than the accepted value for the enthalpy change for this reaction. Some possible sources of error:

* The HCl used in the classroom experiment was from an old stock bottle, and it's possible that some HCl gas escaped over time, making the HCl concentration less than 2.0 M.

* The NaOH used also came from an old stock bottle. If this bottle had been open to the air for a long period of time during previous experiments, carbon dioxide from the air could have reacted with the sodium hydroxide, lowering its concentration.