10 Thermochemistry and Hess's Law

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. What is the definition of the *joule* in terms of the basic SI units?

$J = Kg.m/s^2$

2. In a calorimeter calibration experiment, a sampe of 51.682 g of water at 55.2 °C is added to a calorimeter containing 50.220 g of water at 23.5 °C. After stirring and waiting for the system to equilibrate, the final temperature reached is 36.6 °C. Calculate the calorimeter constant.

$$C = \frac{[m_h \times sp \times (T_h - T_m)] - [m_c \times sp \times (T_m - T_c)]}{T_m - T_c}$$

$$= \frac{[51.203 \times 4.18 \times (55.2 - 37.6) - [49.783 \times 4.18 \times 37.6 - 25.5)}{37.6 - 23.6}$$

$C = 59.06 \ J/C°$

3. Give chemical equations for the reactions that will occur during this Experiment

$NaOH + HCl \longrightarrow NaCl + H_2O$

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Results and Calculations

A. Determination of a Calorimeter Constant.

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	Trial (I)	Trial (II)
Mass (or volume) of cold water	75.00 ml g	g
Temperature of cold water	19.0 °C	°C
Mass (or volume) of hot water	75.00 ml g	g
Temperature of hot water	65.00 °C	°C
Final temperature reached	43.00 °C	°C
Increase of cold water temperature (ΔT)	24.00 °C	°C
Decrease of hot water temperature (ΔT)	22.00 °C	°C
Heat lost by hot water	4.18×75×22 6897 J	J
Heat gained by cold water	7524 J	J
Heat gained by calorimeter water	+ ~~261~~ 627 J	J
Calorimeter constant (C)	26.1 J/°C	J/°C
Average value of calorimeter constant	26.1	J/°C

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B. Heat of Acid/Base Reactions

1. HCl/NaOH

	Trial (I)	Trial (II)
Volume of 1.0 M NaOH used	50.0 mL	mL
Initial temperature of NaOH	20.0 °C	°C
Volume of 1.0 M HCl used	50.0 mL	mL
Initial temperature of HCl	20.0 °C	°C
Final temperature reached	27.0 °C	°C
Total volume of mixture	100.0 mL	mL
Total mass of mixture	100.0 g	g
Temperature change, Δt (final - best HCl/NaOH)	7.0 °C	°C
Average value of the Calorimeter constant (C)	26.1 J/°C	
Heat gained by the solution* (m × s × ΔT solution)	2849 J	J
Heat gained by the calorimeter (C × ΔT)	182.7 J	J
Heat of the reaction	3031.7 J	J
Moles of NaOH reacted	0.05 mol	mol
Moles of HCl reacted	0.05 mol	mol
Moles of water produced	0.05 mol	mol
ΔH ($\frac{q}{n\,mol}$)	-60.63 kJ/mol H_2O	kJ/mol H_2O
Average value of ΔH	-60.63 kJ/mol H_2O	
Literature value of ΔH	-61	- 55.9 kJ/mol H_2O

part A

* Assume density of solution = 1.00 g/mL and specific heat of solution = 4.07 J/g. °C

$d = \frac{m}{v}$

Questions

1. What effect on the calculated calorimeter constant would be observed if the calorimeter cup were made of conducting material (such as metal) rather than plastic foam?

The solution become cold

$\Delta T \downarrow \rightarrow C \uparrow$

2. Why is water typically used as the heat-absorbing liquid in calorimeter?

because the water have high specific heat capacity $Sp = 4.18\ J/g^{\circ}C$

3. Given that:

$$H_2O_{2(\ell)} \rightarrow H_2O_{(\ell)} + \frac{1}{2} O_{2(g)} \qquad \Delta H = -96.0\ kJ/mol$$

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)} \qquad \Delta H = -571.6\ kJ/mol$$

Use Hess's law to calculate ΔH for the reaction:

$$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{2(g)}$$

$H_2O + \frac{1}{2} O_2 \rightarrow H_2O$

$H_2 + \frac{1}{2} O_2 \rightarrow H_2O$

$H_2 + O_2 \rightarrow H_2O$

$\Delta H = \Delta H_1 + \Delta H_2$

$96 - 571.6 = -475.6\ kJ/mol$