A student is given the task of determining the enthalpy of reaction for the reaction between HCl(aq) and NaOH(aq).

Glassware Items	Precision
250 mL Erlenmeyer flasks	± 25 mL
100 mL beakers	± 10 mL
100 mL graduated cylinders	± 0.1 mL

The following materials are available. *2013 practice #2*

1.00 M HCl(aq)	1.00 M NaOH(aq)	distilled water
2.00 M HCl(aq)	2.00 M NaOH(aq)	goggles
insulated cups with covers	gloves	thermometer (± 0.1°C)
stirring rod	lab coat	

The student may select from the glassware listed in the table:

a) The student selects two 100 mL beakers, uses them to measure 50 mL each of 1.00 M HCl(aq) solution and 1.00 M NaOH(aq) solution, and measures an initial temperature of 24.5°C for each solution. Then the student pours the two solutions into an insulated cup, stirs the mixture, covers the cup, and records a maximum temperature of 29.9°C.

(i) Is the experimental design sufficient to determine the enthalpy of reaction to a precision of two significant figures? Justify your answer. *(1pt)*

The beakers can only measure 50 ± 10 mL which will limit the precision to 1 sig fig. DON'T MEASURE W/ BEAKERS

(ii) List two specific changes to the experiment that will allow the student to determine the enthalpy of reaction to a precision of three significant figures. Explain. *(2pts)*

1) use graduated cylinders to measure the solutions (± 0.1 mL = 50.0)

2) use 2.00 M solutions. This will increase ΔT (XX.X°C instead of 5.4°C)

b) A second student is given two solutions, 75.0 mL of 1.00 M HCl and 75.0 mL of 1.00 M NaOH, each at 25.0°C. The student pours the solutions into an insulated cup, stirs the mixture, covers the cup, and records the maximum temperature of the mixture.

(i) The student calculates the amount of heat evolved in the experiment to be 4.1 kJ. So $q_{rxn} = -4.1$ Calculate the student's experimental value for the enthalpy of reaction, in kJ/mol*rxn*. *(2pts)*

$HCl + NaOH \rightarrow H_2O + NaCl$ 1:1 Both are limiting

$0.0750\ L\ HCl \left| \frac{1.00\ mol}{1\ L} \right| = 0.0750\ mol \qquad \frac{-4.1\ kJ}{0.0750\ mol} = -55\ \frac{kJ}{mol}$

(ii) The student assumes that the thermometer and the calorimeter do not absorb energy during the reaction. Does this assumption result in a calculated value of the enthalpy of reaction that is higher than, lower than, or the same as it would have been had the heat capacities of the thermometer and calorimeter been taken into account? Justify your answer. *(2pts)*

The value would be lower (↓) since the thermometer + calorimeter do absorb energy. This lost heat is ignored, so the ΔH_{rxn} would be less than it should have been.

c) A third student calculates a value for the enthalpy of reaction that is significantly higher than the accepted value.

(i) Identify a specific error in procedure made by the student that will result in a calculated value for the enthalpy of reaction that is higher than the accepted value. (Vague statements like "human error" or "incorrect calculations" will not earn credit.) *(1pt)*

The student used 2.00 M solutions instead of 1.00 M.

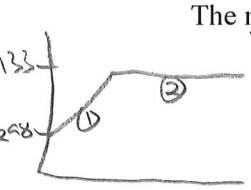
OR The student used more than 75.0 mL of either solution

(ii) Explain how the error that you identified in part (c)(i) leads to a calculated value for the enthalpy of reaction that is higher than the accepted value. *(1pt)*

Larger ΔT would increase q and also the value of ΔH

Aluminum metal can be recycled from scrap metal by melting the metal to evaporate impurities. 2015 #7

a) Calculate the amount of heat needed to purify 1.00 mole of Al originally at 298 K by melting it.
The melting point of Al is 933 K.
The molar heat capacity of Al is 24 J/(mol·K), and the heat of fusion of Al is 10.7 kJ/mol. *(1-3pts)*



1) $q_{solid} = mc\Delta T$
$= (1.00)(24)(635) = 15200\ J = 15.2\ KJ$

2) $q_{fus} = m H_{fus}$
$= (1.00)(10.7) = 10.7\ KJ$

$= \boxed{25.9\ KJ}$

b) The equation for the overall process of extracting Al from Al_2O_3 is shown below. Which requires less energy, recycling existing Al or extracting Al from Al_2O_3 ? Justify your answer with a calculation.

$Al_2O_3(s) \rightarrow 2Al(s) + 3/2\ O_2(g) \quad \Delta H^o = 1675 kJ/mole_{rxn}$ *(2pts)*

$1.00\ mol\ Al \left| \frac{1675\ KJ}{2\ mol\ Al} \right. = \boxed{838\ KJ}$ $25.9\ KJ < 838\ KJ$

★ Recycling Al use much less energy

The heat of combustion of ethanol (C_2H_5OH) is -1367kJ/mole.

a) Determine the ***Calories*** in a 6oz ***(177mL)*** glass of wine if the wine is 10.6% ethanol by mass. Assume the density of the wine is 1.00g/mL and all Calories come from the ethanol.

$$177\ mL\ wine \left| \frac{1.00\ g\ wine}{1\ mL\ wine} \right| \frac{10.6\ g\ EtH}{100\ g\ wine} \left| \frac{1\ mol\ EtH}{46.08\ g\ EtH} \right| \frac{-1367\ KJ\ EtH}{1\ mol\ EtH} \left| \frac{1000\ J}{1\ KJ} \right| \frac{1\ cal}{4.184\ J} \left| \frac{1\ Cal}{1000\ cal} \right.$$

$= \boxed{133\ Cal}$

b) Ethanol has an energy content of about 7Cal/g. Does your calculation in "a" make sense? Justify your answer.

$177\ g\ wine \left| \frac{10.6\ g\ EtH}{100\ g\ wine} \right| \frac{7\ Cal}{1\ g\ EtH} \right. = 131\ Cal$ yes $131 \approx 133$

A student investigates the enthalpy of solution, ΔH_{soln}, for two alkali metal halides, LiCl and NaCl. In addition to the salts, the student has access to a calorimeter, a balance with a precision of ±0.1 g, and a thermometer with a precision of ±0.1°C.
2016 #1 part a

a) To measure ΔH_{soln} for LiCl, the student adds 100.0 g of water initially at 15.0°C to a calorimeter and adds 10.0 g of LiCl(*s*), stirring to dissolve. After the LiCl dissolves completely, the maximum temperature reached by the solution is 35.6°C.

$m = 100.0 + 10.0 = 110.0\ g\ total$

i) Calculate the magnitude of the heat absorbed by the solution during the dissolution process, assuming that the specific heat capacity of the solution is 4.18 J/(g·°C). Include units with your answer. *(2pts)*

$q = mc\Delta T$
$= (110.0)(4.18)(20.6) = 9470\ J = \boxed{9.47\ KJ}$

ii) Determine the value of ΔH_{soln} for LiCl in kJ/mol_{rxn}. *(2pts)*

$10.0\ g\ LiCl \left| \frac{1\ mol\ LiCl}{42.39\ g\ LiCl} \right. = 0.236\ mol\ LiCl$

$\Delta H = \frac{KJ}{mol} = \frac{-9470}{0.236} = \boxed{-40.1\ \frac{KJ}{mol}}$

Answer the following questions about glucose, $C_6H_{12}O_6$, an important biochemical energy source. 2011B #3 a,b

a) Write the empirical formula of glucose. *(1pt)*

$\div 6$ $\quad CH_2O$

In many organisms, glucose is oxidized to carbon dioxide and water, as represented by the following equation.

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

A 2.50 g sample of glucose and an excess of $O_2(g)$ were placed in a calorimeter. After the reaction was initiated and proceeded to completion, the total heat released by the reaction was calculated to be 39.0 kJ.

b) Calculate the value of $\Delta H°$, in kJ mol^{-1}, for the combustion of glucose. *(1pt)*

$2.50\ g\ C_6H_{12}O_6 \left| \frac{1\ mol\ C_6H_{12}O_6}{180.16\ g\ C_6H_{12}O_6} \right| = 0.0139\ mol$ $\quad \Delta H = \frac{-39.0\ KJ}{0.0139\ mol}$

$= \boxed{-2810\ KJ/mol}$

c) When oxygen is not available, glucose can be oxidized by fermentation. In that process, ethanol and carbon dioxide are produced, as represented by the following equation.

$\left(C_6H_{12}O_6(s) \rightarrow 2\ C_2H_5OH(l) + 2\ CO_2(g) \qquad \Delta H° = -68.0\ kJ\ mol^{-1}\ at\ 298\ K \right) \times .5$

Using your answer for part (b) and the information provided above, calculate the value of $\Delta H°$ for the following reaction. $C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(l)$ *(2pts)*

$C_2H_5OH + CO_2 \rightarrow \frac{1}{2} C_6H_{12}O_6 \quad \Delta H = -\frac{1}{2}(-68.0)$

$\frac{1}{2} C_6H_{12}O_6 + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O \quad \Delta H = \frac{1}{2}(-2810)$

$C_2H_5OH + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O \quad \boxed{\Delta H = -1370\ KJ/mol}$

Hydrogen gas burns in air according to the equation: $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$ 2001 #3

0 0 -285.8

a) Calculate the standard enthalpy change, $\Delta H°_{298}$, for the reaction represented by the equation above. (The molar enthalpy of formation, $\Delta H_f°$, for $H_2O(l)$ is -285.8 kJ mol^{-1} at 298 K.) *(1pt)*

Products − Reactants

$\Delta H = [2(-285.8)] - [2(0) + 0] = \boxed{-571.6\ KJ/mol}$

b) Calculate the amount of heat, in kJ, that is released when 10.0 g of $H_2(g)$ is burned in air. *(1pt)*

$10.0\ g\ H_2 \left| \frac{1\ mol\ H_2}{2.02\ g\ H_2} \right| \left| \frac{-571.6\ KJ}{2\ mol\ H_2} \right| = \boxed{-1420\ KJ}$

c) Given that the molar enthalpy of vaporization, $\Delta H°_{vap}$, for $H_2O(l)$ is 44.0 kJ mol^{-1} at 298 K, what is the standard enthalpy change, $\Delta H°_{298}$, for the reaction $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$? *(1pt)*

$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l) \quad \Delta H = -571.6\ KJ$

$2\ H_2O(l) \rightarrow 2\ H_2O(g) \quad \Delta H = 2(44.0\ KJ)$

$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g) \quad \boxed{\Delta H = -483.6\ KJ}$

Answer the following questions about thermodynamics. 2005B #7

Substance	Combustion Reaction	$\Delta H^{o}comb$, at 298 K (kJ mol^{-1})
$H_2(g)$	$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$	−290
$C(s)$	$C(s) + O_2(g) \rightarrow CO_2(g)$	−390
$CH_3OH(l)$		−730

a) Write a balanced chemical equation for the complete combustion of one mole of $CH_3OH(l)$.
Assume products are in their standard states at 298 K. Coefficients do not need to be whole numbers. *(2pts)*

$CH_3OH(l) + \frac{3}{2} O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

b) On the basis of your answer to part (a) and the information in the table, determine the enthalpy change for the reaction

$C(s) + H_2(g) + H_2O(l) \rightarrow CH_3OH(l)$ *(2pts)*

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -390$ kJ

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$ $\Delta H = -290$ kJ

$CO_2(g) + 2H_2O(l) \rightarrow CH_3OH(l) + \frac{3}{2} O_2(g)$ $\Delta H = +730$ kJ

$C(s) + H_2(g) + H_2O(l) \rightarrow CH_3OH(l)$ $\Delta H = 50$ kJ

c) Write the balanced chemical equation that shows the reaction that is used to determine the enthalpy of formation for one mole of $CH_3OH(l)$. *(1pt)*

$C + 2H_2 + \frac{1}{2} O_2 \rightarrow CH_3OH$