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Name:____________________

# Thermodynamics Unit - Quantifying Heat (Group Problems)

1. How much heat, in joules, is required to raise the temperature of 205 g of water from 21.1 °C to 91.4 °C?

Work:

$$q = mC_{water}\Delta T$$

$$q = (205g)(4.184\frac{J}{g^{\circ}C})(91.4^{\circ}C - 21.1^{\circ}C)$$

$$q = 60{,}298\ J$$

2. A constant pressure (coffee cup type) calorimeter having a heat capacity of 472 $J*{}^{o}C^{-1}$ is used to measure the heat evolved when the following aqueous solutions, both initially at 22.6 ºC, are mixed: 100 g of solution containing 6.62 g of lead(II) nitrate, $Pb(NO_3)_2$ , and 100. g of solution containing 6.00 g of sodium iodide, NaI. The final temperature is 24.2ºC. Assume that the specific heat of the mixture is the same as that for water, 4.184 $J*g^{-1}*{}^{o}C^{-1}$ . Calculate the amount of heat evolved in the reaction. Calculate the ΔH of the reaction as written.

– $Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2\ NaNO_3(aq)$

Work:

The mass of the solutions after being mixed is a total of 200 grams.

$$q_{cal} = -\Delta H_{rxn} = C_{cal}\Delta T + mC_{solution}\Delta T$$

$$-\Delta H_{rxn} = (472\frac{J}{^{\circ}C})(24.2^{\circ}C - 22.6^{\circ}C) + (200g)(4.184\frac{J}{g^{\circ}C})(24.2^{\circ}C - 22.6^{\circ}C)$$

$$-\Delta H_{rxn} = 2094\ J$$

$$\Delta H_{rxn} = -2094\ J$$

The reaction raises the temperature of the water and calorimeter, which means that the reaction was exothermic (a negative ΔH).

3. The thermo chemical equation for the combustion of cyclohexane

$$C_6H_{12}(l) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

ΔH = -3920 kJ/mol at 298 K.

What is the change in internal energy for the combustion of 1.00 mol $C_6H_{12}(l)$ at 298K?

Work:

$$\Delta U = q + w$$

$$\Delta U = \Delta H - P\Delta V$$

$$\Delta U = \Delta H - \Delta n_{gas} RT$$

$$\Delta U = (-3920 \frac{kJ}{mol}) - (6mol - 9mol)(0.008314 \frac{kJ}{Kmol})(298K)$$

$$\Delta U = -3912.6 \ kJ$$

4. Compare the energy of combustion of $H_2$ to $CH_4$ using a bomb calorimeter with a heat capacity of 11.3 kJ/ºC. When a 1.50 g sample of methane gas was burned with excess oxygen in the calorimeter, the T increased by 7.3 ºC. When a 1.15 g sample of hydrogen gas was burned with excess oxygen, the temperature increase was 14.3 ºC. Calculate the energy of combustion (per gram) for hydrogen and methane.

First we should write out the combustion reactions for both:

$CH_4 (g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

$H_2 (g) + ½O_2(g) \rightarrow H_2O(l)$

We know that in a bomb calorimeter there is no expansion work. So we will calculate the q value for each reaction.

$Methane:$

$$\Delta U = q_{sys}$$

$$q_{sys} = -q_{cal} = -C_{cal}\Delta T$$

$$q_{sys} = -(11.3 \frac{kJ}{°C})(7.3°C)$$

$$q_{sys} = -83kJ$$

$$\Delta U_{pergram} = \frac{q_{sys}}{mass} = \frac{-83kJ}{1.5g} = -55 \frac{kJ}{g}$$

$Hydrogen:$

$$\Delta U = q_{sys}$$

$$q_{sys} = -q_{cal} = -C_{cal}\Delta T$$

$$q_{sys} = -(11.3 \frac{kJ}{°C})(14.3°C)$$

$$q_{sys} = -162kJ$$

$$\Delta U_{pergram} = \frac{q_{sys}}{mass} = \frac{-162kJ}{1.15g} = -141 \frac{kJ}{g}$$

The combustion reactions are exothermic, which is why the calorimeter experiences a rise in temperature in both cases. The ΔU values are negative because of the negative q values for the reactions.