



## 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?

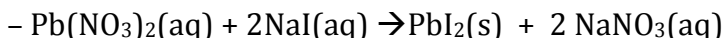
Work:

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

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

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

2. A constant pressure (coffee cup type) calorimeter having a heat capacity of 472 J\*°C<sup>-1</sup> is used to measure the heat evolved when the following aqueous solutions, both initially at 22.6 °C, are mixed: 100 g of solution containing 6.62 g of lead(II) nitrate, Pb(NO<sub>3</sub>)<sub>2</sub>, and 100. g of solution containing 6.00 g of sodium iodide, NaI. The final temperature is 24.2°C. Assume that the specific heat of the mixture is the same as that for water, 4.184 J\*g<sup>-1</sup>\* °C<sup>-1</sup>. Calculate the amount of heat evolved in the reaction. Calculate the ΔH of the reaction as written.



Work:

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

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

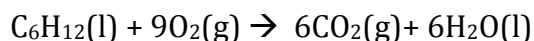
$$-\Delta H_{\text{rxn}} = (472 \frac{\text{J}}{^\circ\text{C}})(24.2^\circ\text{C} - 22.6^\circ\text{C}) + (200\text{g})(4.184 \frac{\text{J}}{\text{g}^\circ\text{C}})(24.2^\circ\text{C} - 22.6^\circ\text{C})$$

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

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

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



$$\Delta H = -3920 \text{ kJ/mol at } 298 \text{ K.}$$

What is the change in internal energy for the combustion of 1.00 mol C<sub>6</sub>H<sub>12</sub>(l) at 298K?



**Work:**

$$\Delta U = q + w$$

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

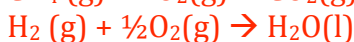
$$\Delta U = \Delta H - \Delta n_{gas}RT$$

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

$$\Delta U = -3912.6 \text{ kJ}$$

4. Compare the energy of combustion of H<sub>2</sub> to CH<sub>4</sub> using a bomb calorimeter with a heat capacity of 11.3 kJ/°C. When a 1.50 g sample of methane gas was burned with excess oxygen in the calorimeter, the T increased by 7.3 °C. When a 1.15 g sample of hydrogen gas was burned with excess oxygen, the temperature increase was 14.3 °C. Calculate the energy of combustion (per gram) for hydrogen and methane.

First we should write out the combustion reactions for both:



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

*Methane :*

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

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

$$q_{sys} = -(11.3 \frac{kJ}{^\circ\text{C}})(7.3^\circ\text{C})$$

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

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

*Hydrogen :*

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

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

$$q_{sys} = -(11.3 \frac{kJ}{^\circ\text{C}})(14.3^\circ\text{C})$$

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

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

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.