Answer on Question #67251, Chemistry / General Chemistry

During an experiment, a student adds 1.05 g of calcium metal to 200.0 mL of 0.75 M HCl. Temperature increase of 17.0 °C for the solution. The solution's final volume is 200.0 mL, the density is 1.00 g/mL, and the specific heat is 4.184 J/(g•°C), calculate the heat of the reaction, ΔH_{rxn}.

$$Ca_{(s)} + 2H^{+}_{aq} \rightarrow Ca^{2+}_{aq} + \ H_2(g)$$

Solution:

The ΔH_{rxn} would be for one mole of Ca reacted or 2 moles of H^+, whichever is the limiting reactant.

$$n(Ca) = \frac{1.05}{40} = 0.026\ (mol)$$
$$n(HCl) = n(H^+) = 0.2 * 0.75 = 0.15\ (mol)$$

Moles of H^+ is more than 2x moles of Ca, so Ca is limiting reactant

Now find ΔH in the experiment:

$$\Delta H = m * c_p * \Delta T$$
$$\Delta H = 200 * 1.00 * 4.184 * 17 = 14225{,}6\ (J)$$

Since the reaction is exothermic,

$$\Delta H_{rxn} = \frac{-\Delta H}{n(Ca)} = \frac{-14225.6}{0.026} = \ -541.9\ (kJ/mol)$$

Answer: -541.9 kJ/mol.

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