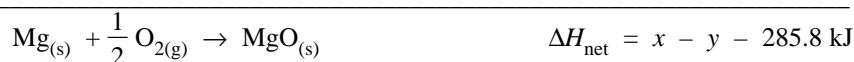
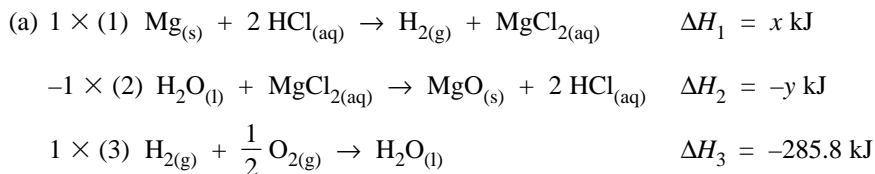


INVESTIGATION 5.4.1 HESS'S LAW

Prediction

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Analysis

(b) Since the temperature of the solution (surroundings) increased, the changes were exothermic and would have to have negative values if written as ΔH values.

(c) (Answers will vary.) Assume a 11.0°C temperature change for 100 mL of acid and 0.500 g of magnesium used:

$$q = mc\Delta T$$

$$= 0.100 \text{ kg} \times 4.18 \text{ kJ/kg}\cdot^\circ\text{C} \times 15.0^\circ\text{C}$$

$$q = 6.27 \text{ kJ}$$

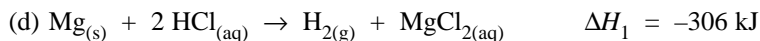
$$n_{\text{Mg}} = \frac{0.500 \text{ g}}{24.3 \text{ g/mol}}$$

$$n_{\text{Mg}} = 0.0205 \text{ mol}$$

$$\Delta H = \frac{q}{n}$$

$$= \frac{6.27 \text{ kJ}}{0.0205 \text{ mol}}$$

$$\Delta H = -306 \text{ kJ/mol Mg } (\Delta H \text{ is given a negative sign because the reaction was exothermic.})$$



(e) (Answers will vary.) Assume a 5.0°C temperature change for 100 mL of acid and 1.00 g of magnesium oxide used:

$$q = mc\Delta T$$

$$= 0.100 \text{ kg} \times 4.18 \text{ kJ/kg}\cdot^\circ\text{C} \times 5.0^\circ\text{C}$$

$$q = 2.1 \text{ kJ}$$

$$n_{\text{MgO}} = \frac{1.00 \text{ g}}{40.3 \text{ g/mol}}$$

$$n_{\text{MgO}} = 0.0248 \text{ mol}$$

$$\Delta H = \frac{q}{n}$$

$$= \frac{2.1 \text{ kJ}}{0.0248 \text{ mol}}$$

$$\Delta H = -84 \text{ kJ/mol MgO } (\Delta H \text{ is given a negative sign because the reaction was exothermic.})$$

$$\text{(f) } \Delta H_{f(\text{H}_2\text{O})} = -285.8 \text{ kJ/mol}$$

$$\Delta H_{\text{net}} = \Delta H_1 - \Delta H_2 - \Delta H_3$$

$$= -306 \text{ kJ} - (-84 \text{ kJ}) - 286 \text{ kJ}$$

$$\Delta H_{\text{net}} = -508 \text{ kJ}$$

The molar enthalpy of combustion of magnesium is -508 kJ .

(g) The temperature measurements had the most uncertainty because of thermometer limitations.

Evaluation

- (h) (i) If heat was lost to the surroundings, the ΔT and ΔH values would be low.
(ii) If MgO was on the surface of the magnesium, less Mg would be present and again the ΔT and ΔH values would be low.
- (i) Temperature and mass measurements may be inaccurate, and heat may be lost to the surroundings.
- (j) % difference = $\frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100\%$
- $$= \frac{|2601.6 \text{ kJ/mol} - (-508 \text{ kJ/mol})|}{2601.6 \text{ kJ/mol}} \times 100\%$$
- % difference = 15.5%

Synthesis

- (k) (Answers will vary.) The accuracy of this method depends on the sophistication of the apparatus. If bomb calorimetry had been used, the values would have been more accurate.

LAB EXERCISE 5.5.1 TESTING ENTHALPIES OF FORMATION

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Prediction

- (a) $\text{CH}_3\text{OH}_{(l)} + 3/2 \text{O}_{2(g)} \rightarrow 1 \text{CO}_{2(g)} + 2 \text{H}_2\text{O}_{(l)}$
- $$\Delta H_{\text{comb (methanol)}} = 1 \Delta H^\circ_{\text{f}(\text{CO}_{2(g)})} + 2 \Delta H^\circ_{\text{f}(\text{H}_2\text{O}_{(l)})} - 1 \Delta H^\circ_{\text{f}(\text{CH}_3\text{OH}_{(l)})} - \frac{3}{2} \Delta H^\circ_{\text{f}(\text{O}_{2(g)})}$$
- $$= 1(-393.5) + 2(-285.8) - 1(-239.1) - \frac{3}{2}(0)$$
- $$\Delta H_{\text{comb (methanol)}} = -726 \text{ kJ}$$
- The molar enthalpy of combustion of methanol is -726 kJ .

Analysis

- (b) $q_{\text{water}} = c \times \Delta T$
- $$= 10.9 \text{ kJ/}^\circ\text{C} \times (27.9^\circ\text{C} - 20.4^\circ\text{C})$$
- $$q_{\text{water}} = 81.7(5) \text{ kJ}$$
- $$n_{\text{methanol}} = 4.38 \text{ g} \times \frac{1 \text{ mol}}{32.0 \text{ g}}$$
- $$n_{\text{methanol}} = 0.137 \text{ mol}$$
- $$\Delta H_{\text{comb (methanol)}} = \frac{q_{\text{water}}}{n}$$
- $$= \frac{81.7(5) \text{ kJ}}{0.137 \text{ mol}}$$
- $$\Delta H_{\text{comb (methanol)}} = 597 \text{ kJ/mol}$$

Because the reaction is exothermic, the molar heat of combustion of methanol is -597 kJ/mol .

Evaluation

- (c) % difference = $\frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100\%$
- $$= \frac{|726 \text{ kJ/mol} - (-597 \text{ kJ/mol})|}{726 \text{ kJ/mol}} \times 100\%$$
- % difference = 18%

This is a reasonable result, given relatively simple apparatus.