# Thermochemistry Review - Answers

- Using the data below, calculate the amount of energy required to melt a 16.2 g aluminum pop can. Assume that the aluminum is at an initial temperature of 20°C.
  - Molar heat of fusion of Al(s) = 10.7 kJ/mol

  - Melting point of Al(s) = 660°C
     Heat capacity of Al(s) = 0.900 J/g°C

mol = 16.2 q - 26.98154 q/mol = 0.600 mol ΔH = (20°C → 660°C) + (melt)  $\Delta H = (mc\Delta T) + (n\Delta H fus)$  $\Delta H = (16.2 \text{ g})(0.900 \text{ J/g}^{\circ}\text{C})(640^{\circ}\text{C}) + (0.600 \text{ mol})(10.7 \text{ kJ/mol})$  $\Delta H = 9331.2 \text{ J} \cdot 6.42 \text{ kJ}$  $\Delta H = 9.33 \text{ kJ} + 6.42 \text{ kJ}$ 

Given the thermochemical equations:

AH = 15.8 kJ

 $C(s) \cdot O_t(q) \rightarrow CO_t(q)$  $\Delta H = -394 \text{ kJ}$  $\Delta H = -286 \text{ kJ}$  $H_r(a) \cdot \{O_r(a) \rightarrow H_rO(l)$ C(s) • 2 H<sub>2</sub>(q) • ↑ O<sub>2</sub>(q) → CH<sub>2</sub>OH(I) ΔH = -239 kJ

Calculate the  $\Delta H$  for the reaction:  $CH_2OH(1) - {}^3/{}_2O_2(q) \rightarrow CO_2(q) + 2H_2O(1)$ 

 $CH_2OH(1) \rightarrow C(s) + 2 H_2(q) + \frac{1}{2} O_2(q)$   $\Delta H = (-1)(-239 kJ)$  $C(s) \cdot O_2(q) \rightarrow CO_2(q)$  $\Delta H = (1)(-394 \text{ kJ})$ 2 H₂(a) + O₂(a) → 2 H₂O(l)  $\Delta H = (2)(-286 \text{ kJ})$ 

 $CH_3OH(1) + {}^3/_2 O_2(q) \rightarrow CO_2(q) + 2 H_2O(1)$   $\Delta H = -727 \text{ kJ}$ 

# **Thermochemistry Review Answers**

**ACT** 

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