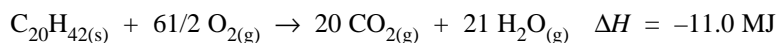


$$\begin{aligned}
 q_{\text{copper}} &= mc\Delta T \\
 &= 50.0 \text{ g} \times 0.385 \text{ J/(g}\cdot^{\circ}\text{C)} \times (76.0 - 21.0)^{\circ}\text{C} \\
 q_{\text{copper}} &= 1.06 \text{ kJ} \\
 n\Delta H_{\text{reaction}} &= q_{\text{total}} \\
 &= q_{\text{water}} + q_{\text{copper}} \\
 n\Delta H_{\text{reaction}} &= 47.0(4) \text{ kJ [Digit in parentheses will be lost in rounding.]} \\
 m_{\text{eicosane}} &= 8.567 - 7.357 \text{ g} \\
 m_{\text{eicosane}} &= 1.21 \text{ g} \\
 M_{\text{eicosane (C}_{20}\text{H}_{42})} &= 282 \text{ g/mol} \\
 n_{\text{eicosane}} &= 1.21 \text{ g} \times \frac{1 \text{ mol}}{282 \text{ g}} \\
 n_{\text{eicosane}} &= 4.29 \times 10^{-3} \text{ mol} \\
 \Delta H_{\text{comb}} &= \frac{q_{\text{total}}}{n} \\
 &= \frac{47.0(4) \text{ kJ}}{4.29 \times 10^{-3} \text{ mol}} \\
 \Delta H_{\text{comb}} &= 1.10 \times 10^4 \text{ kJ/mol C}_{20}\text{H}_{42}
 \end{aligned}$$

Because the reaction is exothermic,  $\Delta H_{\text{reaction}}$  is  $-11.0 \text{ MJ/mol}$ .

(b) The reaction was exothermic, because heat was released to the surroundings and the temperature increased.

(c)  $\text{C}_{20}\text{H}_{42(\text{s})} + 61/2 \text{ O}_{2(\text{g})} \rightarrow 20 \text{ CO}_{2(\text{g})} + 21 \text{ H}_2\text{O}_{(\text{g})} + 11.0 \text{ MJ}$



#### Evaluation

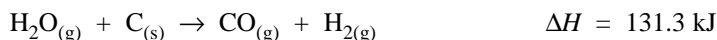
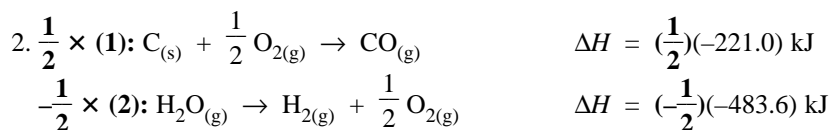
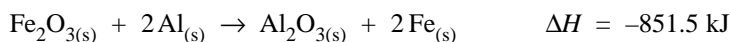
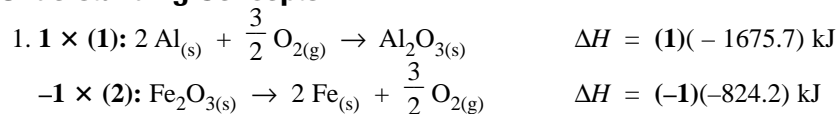
$$\begin{aligned}
 \text{(d) Percentage error} &= \frac{(13.3 - 11.0)}{13.3} \times 100 \\
 &= 17\%
 \end{aligned}$$

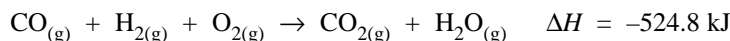
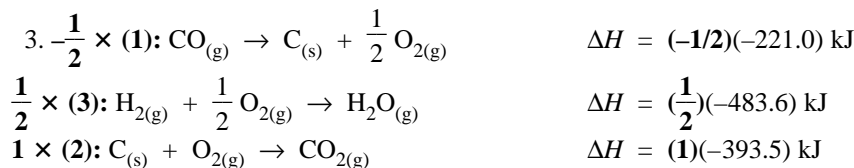
## 5.4 HESS'S LAW OF ADDITIVITY OF REACTION ENTHALPIES

### PRACTICE

(Page 326)

#### Understanding Concepts

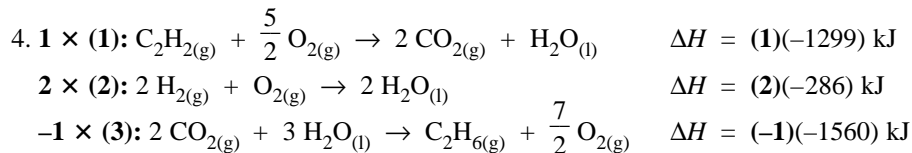




## PRACTICE

(Page 329)

### Understanding Concepts



Since the reaction is written for one mole of ethyne, the molar enthalpy of combustion of ethyne,  $\Delta H_{\text{comb}}$ , is  $-311 \text{ kJ/mol}$ .

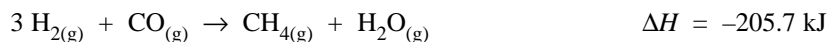
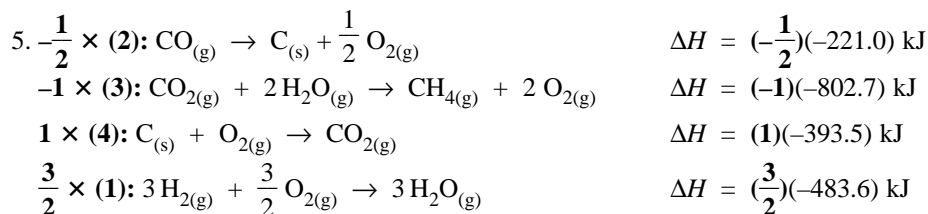
$$\text{amount of ethyne, } n = 200 \text{ g ethyne} \times \frac{1 \text{ mol}}{26.0 \text{ g}}$$

$$n = 7.69 \text{ mol}$$

$$\Delta H = n\Delta H_{\text{comb}}$$

$$= 7.69 \text{ mol} \times (-311 \text{ kJ/mol})$$

$$\Delta H = -2.39 \times 10^3 \text{ kJ}$$



Since the reaction is written for one mole of carbon monoxide, the molar enthalpy of combustion,  $\Delta H_{\text{comb}}$ , is  $-205.7 \text{ kJ/mol CO}$ .

$$\text{amount of CO, } n = 300 \text{ g} \times 1 \text{ mol} / 28.0 \text{ g}$$

$$n = 10.7 \text{ mol}$$

$$\Delta H = n\Delta H_{\text{comb}}$$

$$= 10.7 \text{ mol} \times (-205.7 \text{ kJ/mol})$$

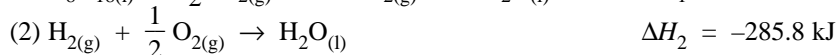
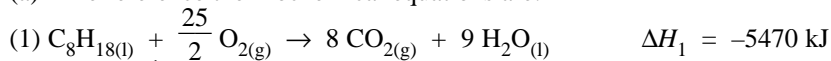
$$\Delta H = -2.20 \times 10^3 \text{ kJ}$$

## SECTION 5.4 QUESTIONS

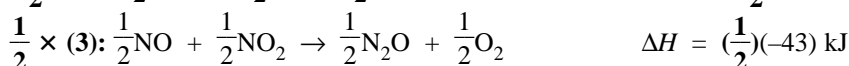
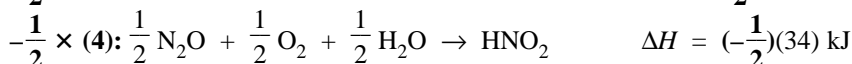
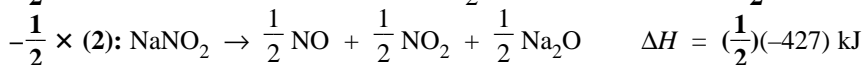
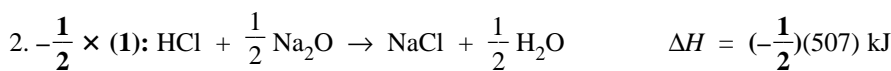
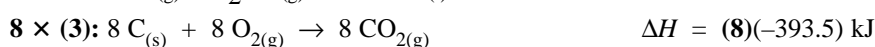
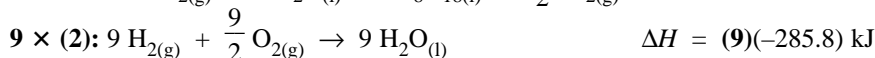
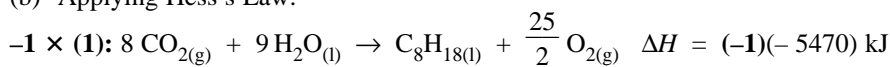
(Page 330)

### Understanding Concepts

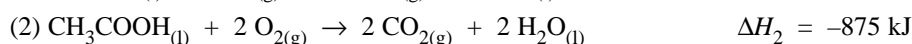
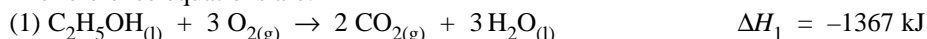
1. (a) The reference thermochemical equations are:



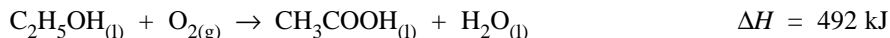
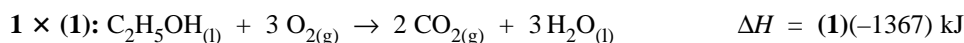
(b) Applying Hess's Law:



3. The reference equations are:

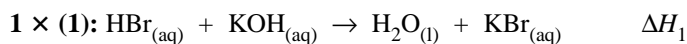


Applying Hess's Law:



### Applying Inquiry Skills

4. (a) When reference equations (1) and (2) are added together, the result is reference equation (3).



(b) In all three experiments, assume 200.0 mL of solution is 200.0 g water.

**Experiment 1:**

$$q_{\text{water}} = mc\Delta T$$

$$= 200.0 \text{ g} \times 4.18 \text{ J/g}\cdot^{\circ}\text{C} \times (22.5 - 20.0)^{\circ}\text{C}$$

$$q_{\text{water}} = 2.09 \text{ kJ}$$

$$n\Delta H_1 = q_{\text{water}}$$

$$= 2.09 \text{ kJ}$$

$$\text{amount of KOH, } n = CV$$

$$= 1.00 \text{ mol/L} \times 0.100 \text{ L}$$

$$n = 0.100 \text{ mol}$$

$$\Delta H_1 = \frac{q}{n}$$

$$= \frac{2.09 \text{ kJ}}{0.100 \text{ mol}}$$

$$\Delta H_1 = 20.9 \text{ kJ/mol KOH}$$

Because the reaction is exothermic and is written for one mole of KOH,  $\Delta H_1$  is  $-20.9 \text{ kJ}$ .

**Experiment 2:**

$$q_{\text{water}} = mc\Delta T$$

$$= 200.0 \text{ g} \times 4.18 \text{ J/g}\cdot^{\circ}\text{C} \times (24.1 - 20.0)^{\circ}\text{C}$$

$$q_{\text{water}} = 3.43 \text{ kJ}$$

$$n\Delta H_2 = q_{\text{water}}$$

$$= 3.43 \text{ kJ}$$

$$\text{molar mass of KOH, } M = 56.1 \text{ g/mol}$$

$$\text{amount of KOH, } n = 5.61 \text{ g} \times \frac{1 \text{ mol}}{56.1 \text{ g}}$$

$$n = 0.100 \text{ mol}$$

$$\Delta H_2 = \frac{q}{n}$$

$$= \frac{3.43 \text{ kJ}}{0.100 \text{ mol}}$$

$$\Delta H_2 = 34.3 \text{ kJ/mol KOH}$$

Because the reaction is exothermic and is written for one mole of KOH,  $\Delta H_2$  is  $-34.3 \text{ kJ}$ .

**Experiment 3:**

$$q_{\text{water}} = mc\Delta T$$

$$= 200.0 \text{ g} \times 4.18 \text{ J/g}\cdot^{\circ}\text{C} \times (26.7 - 20.0)^{\circ}\text{C}$$

$$q_{\text{water}} = 5.60 \text{ kJ}$$

$$n\Delta H_{\text{reaction}} = q_{\text{water}}$$

$$= 5.60 \text{ kJ}$$

$$\text{amount of KOH, } n = 5.61 \text{ g} \times \frac{1 \text{ mol}}{56.1 \text{ g}}$$

$$n = 0.100 \text{ mol}$$

$$\Delta H_3 = \frac{q}{n}$$

$$= \frac{5.60 \text{ kJ}}{0.100 \text{ mol}}$$

$$\Delta H_3 = 56.0 \text{ kJ/mol KOH}$$

Because the reaction is exothermic and is written for one mole of KOH,  $\Delta H_3$  is  $-56.0 \text{ kJ/mol}$ .

$$\begin{aligned}
 \text{(c) Experimental percentage error} &= \frac{|(\Delta H_1 + \Delta H_2) - \Delta H_3|}{\Delta H_3} \times 100\% \\
 &= \frac{|(20.9 + 34.3) - 56.0|}{56.0} \times 100\% \\
 &= 1.4\%
 \end{aligned}$$

## 5.5 STANDARD ENTHALPIES OF FORMATION

### PRACTICE

(Page 332)

#### Understanding Concepts

- $6 \text{C}_{(s)} + 3 \text{H}_{2(g)} \rightarrow \text{C}_6\text{H}_{6(l)}$
  - $\text{K}_{(s)} + \frac{1}{2} \text{Br}_{2(l)} + \frac{3}{2} \text{O}_{2(g)} \rightarrow \text{KBrO}_{3(s)}$
  - $6 \text{C}_{(s)} + 6 \text{H}_{2(g)} + 3 \text{O}_{2(g)} \rightarrow \text{C}_6\text{H}_{12}\text{O}_{6(s)}$
  - $\text{Mg}_{(s)} + \text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow \text{Mg}(\text{OH})_{2(s)}$

### PRACTICE

(Page 335)

#### Understanding Concepts

- $\text{C}_5\text{H}_{12(l)} + 8 \text{O}_{2(g)} \rightarrow 5 \text{CO}_{2(g)} + 6 \text{H}_2\text{O}_{(l)}$ 

$$\Delta H^\circ_{f(\text{C}_5\text{H}_{12(l)})} = -173.5 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{O}_{2(g)})} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{CO}_{2(g)})} = -393.5 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{H}_2\text{O}_{(l)})} = -285.8 \text{ kJ/mol}$$

$$\Delta H = 5 \Delta H^\circ_{f(\text{CO}_{2(g)})} + 6 \Delta H^\circ_{f(\text{H}_2\text{O}_{(l)})} - 1 \Delta H^\circ_{f(\text{C}_5\text{H}_{12(l)})} - 8 \Delta H^\circ_{f(\text{O}_{2(g)})}$$

$$= 5(-393.5) + 6(-285.8) - 1(-173.5) - 8(0)$$

$$\Delta H = -3509 \text{ kJ}$$
  - $\text{Fe}_2\text{O}_{3(s)} + 3 \text{CO}_{(g)} \rightarrow 2 \text{Fe}_{(s)} + 3 \text{CO}_{2(g)}$ 

$$\Delta H^\circ_{f(\text{Fe}_2\text{O}_{3(s)})} = -824.2 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{CO}_{(g)})} = -110.5 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{Fe}_{(s)})} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{CO}_{2(g)})} = -393.5 \text{ kJ/mol}$$

$$\Delta H = 2 \Delta H^\circ_{f(\text{Fe}_{(s)})} + 3 \Delta H^\circ_{f(\text{CO}_{2(g)})} - 1 \Delta H^\circ_{f(\text{Fe}_2\text{O}_{3(s)})} - 3 \Delta H^\circ_{f(\text{CO}_{(g)})}$$

$$= 2(0) + 3(-393.5) - 1(-824.2) - 3(-110.5)$$

$$\Delta H = -24.8 \text{ kJ}$$
- $\text{C}_6\text{H}_{12(l)} + 9 \text{O}_{2(g)} \rightarrow 6 \text{CO}_{2(g)} + 6 \text{H}_2\text{O}_{(l)}$ 

$$\Delta H^\circ_{\text{comb}(\text{C}_6\text{H}_{12(l)})} = -3824 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{O}_{2(g)})} = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{CO}_{2(g)})} = -393.5 \text{ kJ/mol}$$

$$\Delta H^\circ_{f(\text{H}_2\text{O}_{(l)})} = -285.8 \text{ kJ/mol}$$