

CHEM 1215 Unit 1 Practice Problems - Solutions

$$1. \text{ kJ} = 1.525 \text{ kcal} \times \frac{10^3 \text{ cal}}{1 \text{ kcal}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{10^3 \text{ J}} = \boxed{6.381 \text{ kJ}}$$

$$2. \text{ Given: } q = +125 \text{ J}$$

$$w = -0.110 \text{ kJ} \times \frac{10^3 \text{ J}}{\text{kJ}} = -110. \text{ J}$$

$$\text{ Find: } \Delta E = q + w = 125 \text{ J} + -110. \text{ J} = \boxed{15 \text{ J}}$$

$$3. \text{ Given: } \left. \begin{array}{l} V_f = 4.50 \text{ L} \\ V_i = 1.50 \text{ L} \end{array} \right\} \Delta V = 3.00 \text{ L}$$

$$P = 1.25 \text{ atm}$$

$$\Delta H = 1.515 \text{ kJ}$$

$$\text{ Find: } \Delta E = \Delta H - P\Delta V$$

$$\Delta E = 1.515 \text{ kJ} - \left[(1.25 \text{ atm})(3.00 \text{ L}) \left(\frac{101.3 \text{ J}}{1 \text{ atm} \cdot \text{L}} \right) \left(\frac{1 \text{ kJ}}{10^3 \text{ J}} \right) \right]$$

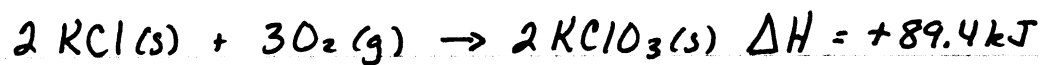
$$= 1.515 \text{ kJ} - 0.379875 \text{ kJ} = 1.135125 \text{ kJ} \Rightarrow \boxed{1.135 \text{ kJ}}$$

$$4. \text{ kJ} = 2.5 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{-89.4 \text{ kJ}}{2 \text{ mol KClO}_3}$$

$$= -0.9187 \text{ kJ}$$

$$= \boxed{-0.91 \text{ kJ}}$$

5. Note: This problem asks about the Formation of KClO_3 - the reverse of the equation given in question 4. Write the thermochemical eq'n for the reverse reaction:



So,

$$\Delta H = 7.4 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{+89.4 \text{ kJ}}{2 \text{ mol KClO}_3}$$

$$= 2.6991 \text{ kJ} \Rightarrow \boxed{2.7 \text{ kJ}}$$

6. Given: $C_s = 0.108 \text{ cal/gK}$
 0.250 mol Fe
 $T_f = 1535^\circ\text{C}$
 $T_i = 25^\circ\text{C}$ } $\Delta T = 1510. \text{ K}$

Find: q (kJ)

$$q = C_s \times \text{mass} \times \Delta T$$

$$q = 0.108 \frac{\text{cal}}{\text{g} \cdot \text{K}} \times 0.250 \text{ mol} \times \frac{55.85 \text{ g}}{1 \text{ mol}} \times 1510. \text{ K} \times \frac{4.184 \text{ J}}{1 \text{ cal}}$$

$$\times \frac{1 \text{ kJ}}{10^3 \text{ J}} = \boxed{9.53 \text{ kJ}}$$

7. Given: $q = 1.376 \text{ kJ}$
 $\text{mass} = 55.0 \text{ g}$
 $T_f = 87.5^\circ\text{C}$
 $T_i = 22.5^\circ\text{C}$ } $\Delta T = 65.0 \text{ K}$

Find: C_s (J/gK)

$$C_s = \frac{q}{\text{mass} \cdot \Delta T} = \frac{1.376 \text{ kJ}}{55.0 \text{ g} \times 65.0 \text{ K}} \times \frac{10^3 \text{ J}}{1 \text{ kJ}}$$

$$= \boxed{0.385 \text{ J/g}\cdot\text{K}}$$

8. Given: $\text{mass metal} = 26.4 \text{ g}$ $\text{mass H}_2\text{O} = 25.0 \text{ g}$
 $T_f(\text{metal}) = 27.0^\circ\text{C}$ $T_f(\text{H}_2\text{O}) = 27.0^\circ\text{C}$
 $T_i(\text{metal}) = 99.0^\circ\text{C}$ $T_i(\text{H}_2\text{O}) = 20.0^\circ\text{C}$
 $\Delta T(\text{metal}) = -72.0^\circ\text{C}$ $\Delta T(\text{H}_2\text{O}) = 7.0^\circ\text{C}$
 $M(\text{metal}) = 63.5 \text{ g/mol}$ $C_s(\text{H}_2\text{O}) = 4.18 \text{ J/g}\cdot\text{K}$

Find: C_s & C_m

$$q_{\text{H}_2\text{O}} = C_s(\text{H}_2\text{O}) \times \text{mass H}_2\text{O} \times \Delta T_{\text{H}_2\text{O}}$$

$$= 4.18 \frac{\text{J}}{\text{g}\cdot\text{K}} \times 25.0 \text{ g} \times 7.0 \text{ K} = 731.5 \text{ J}$$

$$q_{\text{metal}} = -q_{\text{H}_2\text{O}} = -731.5 \text{ J}$$

$$C_s(\text{metal}) = \frac{-731.5 \text{ J}}{26.4 \text{ g} \cdot -72.0 \text{ K}} = \boxed{0.38 \text{ J/g}\cdot\text{K}}$$

8. (cont)

Since $M_{\text{metal}} = 63.5 \text{ g/mol}$:

$$C_m(\text{metal}) = 0.3848 \frac{\text{J}}{\text{g}\cdot\text{K}} \times \frac{63.5 \text{ g}}{\text{mol}} = 24.43$$
$$= \boxed{24 \text{ J/mol}\cdot\text{K}}$$

(Note: $0.3848 \frac{\text{J}}{\text{g}\cdot\text{K}}$ was used to

minimize rounding error.)

9. Given: mass solute = 5.0g KBr } mass soln = 35.0g
mass H₂O = 30.0g }
 $T_f = 15.3^\circ\text{C}$ } $\Delta T_{\text{soln}} = -5.7\text{K}$
 $T_i = 21.0^\circ\text{C}$ }
 $C_s(\text{soln}) = 4.18 \text{ J/gK}$

Find: ΔH_{soln} (kJ/mol)

$$q_{\text{soln}} = \frac{4.18 \text{ J}}{\text{g}\cdot\text{K}} \times 35.0 \text{ g} \times -5.7 \text{ K} = -833.9 \text{ J}$$

$$q_{\text{KBr}} = -q_{\text{soln}} = -(-833.9 \text{ J}) = +833.9 \text{ J}$$

$$\Delta H_{\text{soln}} = \frac{q_{\text{KBr}}}{\text{mol KBr}} = \frac{+833.9 \text{ J}}{5.0 \text{ g KBr}} \times \frac{119.00 \text{ g KBr}}{1 \text{ mol KBr}} \times \frac{1 \text{ kJ}}{10^3 \text{ J}}$$

$$= 19.8468 \frac{\text{kJ}}{\text{mol}} \Rightarrow$$

$$\boxed{20. \text{ kJ/mol}}$$

10. Given: mass = 1.276g C₃H₈
 $C_{cal} = 7.850 \text{ kJ/K}$
 $T_f = 29.21^\circ\text{C}$
 $T_i = 21.01^\circ\text{C}$ } $\Delta T = 8.20^\circ\text{C} = 8.20 \text{ K}$

Find: ΔH_{comb} (kJ/mol)

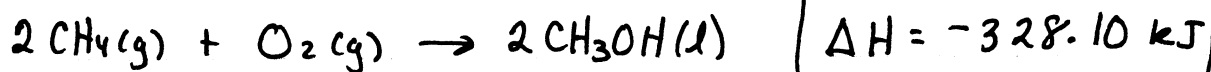
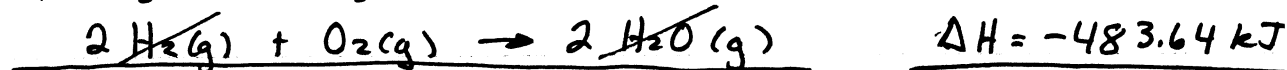
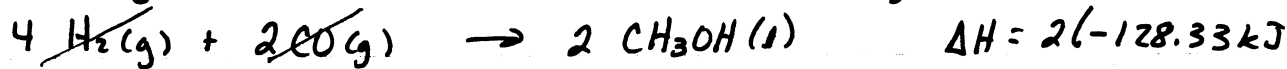
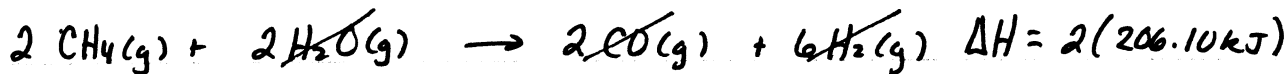
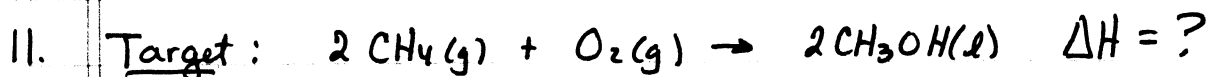
$$q_{cal} = C_{cal} \times \Delta T \quad (\text{notice: NO mass term!})$$

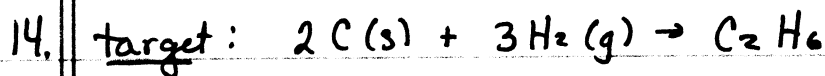
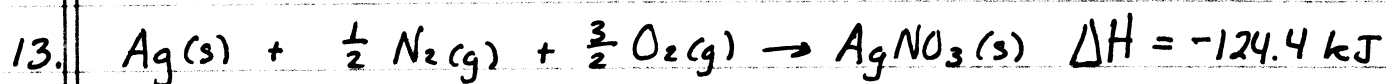
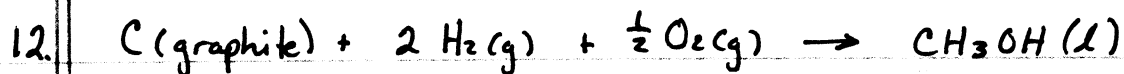
$$= \frac{7.850 \text{ kJ}}{\text{K}} \times 8.20 \text{ K} = 64.37 \text{ kJ}$$

$$q_{C_3H_8} = -q_{cal} = -64.37 \text{ kJ}$$

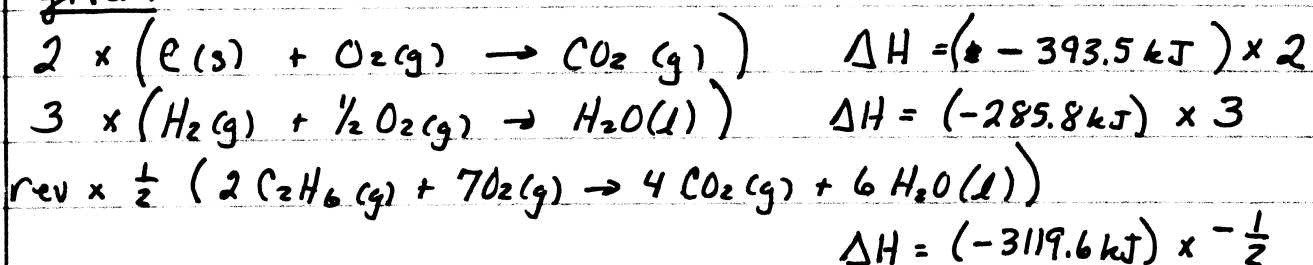
$$\Delta H_{comb} = \frac{-64.37 \text{ kJ}}{1.276 \text{ g C}_3\text{H}_8} \times \frac{44.09 \text{ g C}_3\text{H}_8}{1 \text{ mol C}_3\text{H}_8}$$

$$= -2224.195 \text{ kJ/mol} \Rightarrow \boxed{-2220 \text{ kJ/mol}}$$

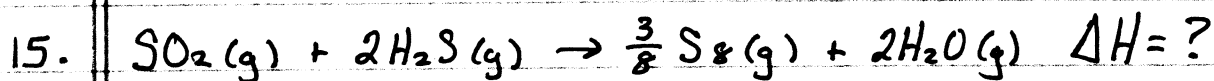
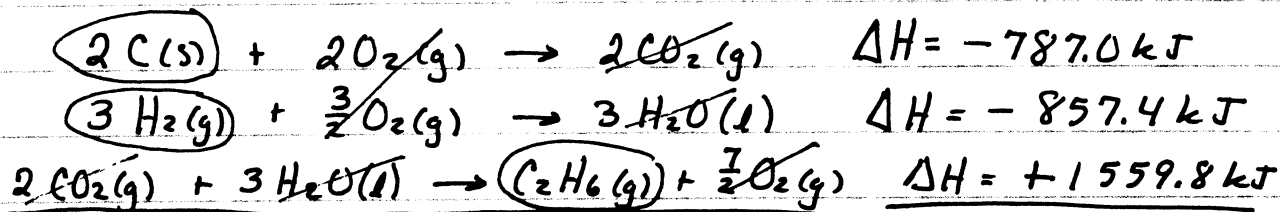




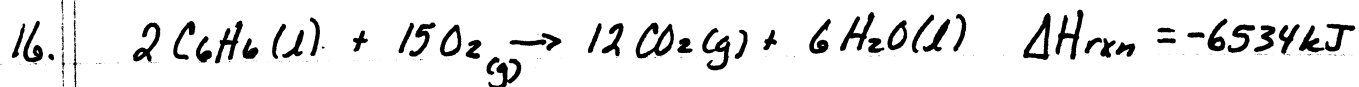
given:



So:



$$\begin{aligned} \Delta H_{\text{rxn}}^{\circ} &= \left[\left(\frac{3}{8} \text{mol} \right) \left(\frac{102.3 \text{ kJ}}{\text{mol}} \right) + \left(2 \text{mol} \right) \left(\frac{-241.82 \text{ kJ}}{\text{mol}} \right) \right] - \\ &\quad \left[\left(1 \text{mol} \right) \left(\frac{-296.9 \text{ kJ}}{\text{mol}} \right) + \left(2 \text{mol} \right) \left(\frac{-20.17 \text{ kJ}}{\text{mol}} \right) \right] \\ &= (38.36 \text{ kJ} + -483.64 \text{ kJ}) - (-296.9 \text{ kJ} + -40.34 \text{ kJ}) \\ &= -108.04 \text{ kJ} \Rightarrow -108.0 \text{ kJ} \end{aligned}$$



$$\Delta H_f^\circ (\text{CO}_2(\text{g})) = -393.5 \text{ kJ/mol}$$

$$\Delta H_f^\circ (\text{H}_2\text{O}(\text{l})) = -285.83 \text{ kJ/mol}$$

$$\Delta H_f^\circ (\text{O}_2(\text{g})) = 0$$

$$\Delta H_f^\circ (\text{C}_6\text{H}_6(\text{l})) = ?$$

$$-6534 \text{ kJ} = \left[(12 \text{ mol}) \left(\frac{-393.5 \text{ kJ}}{\text{mol}} \right) + (6 \text{ mol}) \left(\frac{-285.83 \text{ kJ}}{\text{mol}} \right) \right] - \left[(2 \text{ mol C}_6\text{H}_6) \left(\frac{\Delta H_f^\circ}{\text{C}_6\text{H}_6} \right) \right]$$

$$-6534 \text{ kJ} = -6436.98 \text{ kJ} - (2 \text{ mol C}_6\text{H}_6) (\Delta H_f^\circ (\text{C}_6\text{H}_6))$$

$$(2 \text{ mol C}_6\text{H}_6) (\Delta H_f^\circ \text{C}_6\text{H}_6) = -6436.98 \text{ kJ} + 6534 \text{ kJ}$$

$$\frac{(2 \text{ mol C}_6\text{H}_6) (\Delta H_f^\circ \text{C}_6\text{H}_6)}{2 \text{ mol C}_6\text{H}_6} = \frac{97.02 \text{ kJ}}{2 \text{ mol C}_6\text{H}_6}$$

$$\Delta H_f^\circ (\text{C}_6\text{H}_6) = \frac{48.51 \text{ kJ}}{\text{mol C}_6\text{H}_6} = \boxed{49 \text{ kJ/mol}}$$

