

Ch. 6 Study Questions (*) I assumed 3 s.f. for each problem

1. a) $\frac{500 \text{ J}}{10.9^\circ\text{C}} = \boxed{\frac{45.9 \text{ J}}{^\circ\text{C}}}$ or $\frac{45.9 \text{ J}}{\text{K}}$

b) $\frac{500 \text{ J}}{(500 \text{ g})(10.9^\circ\text{C})} = \boxed{\frac{0.917 \text{ J}}{\text{g}^\circ\text{C}}}$ or $\frac{0.917 \text{ J}}{\text{g}\cdot\text{K}}$

2. $q = (1000 \text{ g}) \left(\frac{0.451 \text{ J}}{^\circ\text{C}} \right) (975^\circ\text{C}) = 439,725 \text{ J} = \boxed{4.40 \times 10^5 \text{ J}}$

3. $-q_{\text{copper}} = +q_{\text{H}_2\text{O}}$

$-(40 \text{ g}) \left(\frac{0.385 \text{ J}}{\text{g}^\circ\text{C}} \right) (T_f - 100^\circ\text{C}) = (100 \text{ g}) \left(\frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \right) (T_f - 25^\circ\text{C})$

$-15.4 T_f + 1540 = 418.4 T_f - 10460$

$T_f = \boxed{27.7^\circ\text{C}}$

4. $27.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{6.009 \text{ kJ}}{1 \text{ mol}} = \boxed{9.01 \text{ kJ}}$

5. Heat to melt: $27.0 \text{ g @ } 0^\circ\text{C} \rightarrow 9.01 \text{ kJ or } 9010 \text{ J}$

$+ (q_{\text{fus}} + q_{\text{ice}}) = -q_{\text{H}_2\text{O}}$

$9010 \text{ J} + (27.0 \text{ g}) \left(\frac{4.184 \text{ J}}{^\circ\text{C}\cdot\text{g}} \right) (T_f - 0^\circ\text{C}) = - (123 \text{ g}) \left(\frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \right) (T_f - 100^\circ\text{C})$

$9010 \text{ J} + 112.968 T_f = -514.632 T_f + 51463.2$

$T_f = \boxed{67.6^\circ\text{C}}$

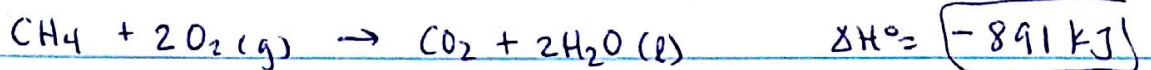
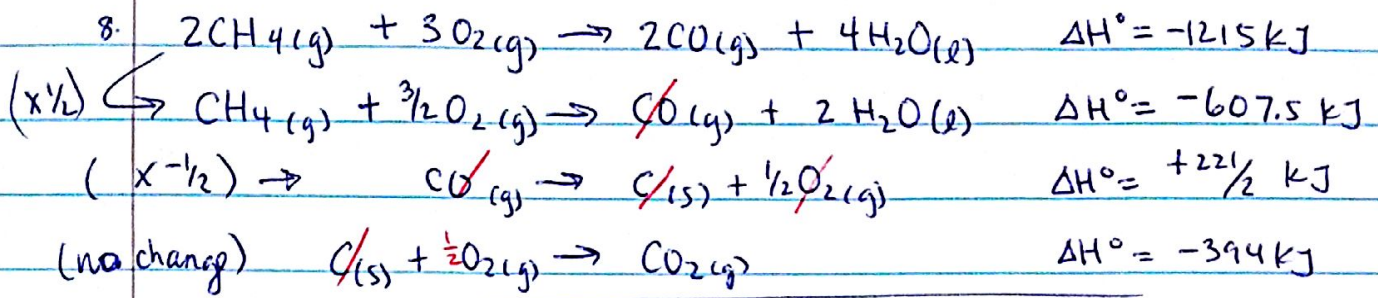
6. $10.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{6.009 \text{ kJ}}{1 \text{ mol}} = 3.50 \text{ kJ} = 3,500 \text{ J}$

$q_{\text{metal}} = m c \Delta T$

$-3,500 \text{ J} = (50 \text{ g})(c)(0^\circ\text{C} - 100^\circ\text{C})$

$c = \boxed{\frac{0.700 \text{ J}}{\text{g}^\circ\text{C}}}$

$$7. \quad 11 \text{ kg C}_3\text{H}_8 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol C}_3\text{H}_8}{44.094 \text{ g}} \times \frac{-2220 \text{ kJ}}{1 \text{ mol C}_3\text{H}_8} = \boxed{-50,300 \text{ kJ}}$$



$$9. \quad \Delta H_{\text{rxn}} = [\Delta H_f(\text{CO}_2) + 2(\Delta H_f \text{H}_2\text{O})] - [\Delta H_f \text{CH}_4 + 2\Delta H_f \text{O}_2]$$

$$-891 = -394 + 2(-286) - X + 0$$

$$X = \Delta H_f^\circ \text{CH}_4 = \boxed{-75 \text{ kJ/mol}}$$

$$10. \quad \Delta H_{\text{rxn}}^\circ = 2(\Delta H_f \text{NO}_2) + 3(\Delta H_f \text{H}_2\text{O}) - 2(\Delta H_f \text{NH}_3) - 0$$

$$-698 = 2(33) + 3(-286) - 2X$$

$$2X = -94$$

$$X = \boxed{-47 \text{ kJ/mol}} = \Delta H_f^\circ \text{NH}_3$$

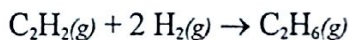
$$11. \quad q_{\text{soln}} = -q_{\text{rxn}}$$

$$= -m_{\text{H}_2\text{O}} c \Delta T$$

$$= -(100 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(-22.4^\circ\text{C})$$

$$= 9372.16 \text{ J}$$

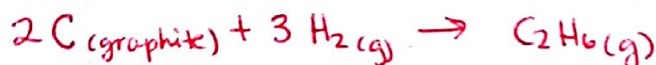
$$\Delta H_{\text{soln}} = \frac{9.37216 \text{ kJ}}{(40 \text{ g} \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.052 \text{ g}})} = \boxed{\frac{18.8 \text{ kJ}}{\text{mol}}}$$

6 • Thermochemistry**PRACTICE FRQ**

Information about the substances involved in the reaction represented above is summarized in the following tables.

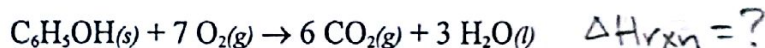
Substance	ΔH_f° (kJ/mol)
$\text{C}_2\text{H}_2(\text{g})$	226.7
$\text{C}_2\text{H}_6(\text{g})$	-84.7

- (a) Write the equation for the heat of formation of $\text{C}_2\text{H}_6(\text{g})$



- (b) Use the above information to determine the enthalpy of reaction for the equation given.

$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= \Delta H_f^\circ \text{C}_2\text{H}_6 - (\Delta H_f^\circ \text{C}_2\text{H}_2 + 2\Delta H_f^\circ \text{H}_2) \\ &= -84.7 - (+226.7 + 0) \\ &= \boxed{-311.4 \frac{\text{kJ}}{\text{mol}}} \end{aligned}$$



When a 2.000-gram sample of pure phenol, $\text{C}_6\text{H}_5\text{OH}(\text{s})$, is completely burned according to the equation above, 64.98 kilojoules of heat is released. Use the information in the table below to answer the questions that follow.

Substance	Standard Heat of Formation, ΔH_f° at 25°C (kJ/mol)
$\text{CO}_2(\text{g})$	-393.5
$\text{H}_2\text{O}(\text{l})$	-285.85
$\text{C}_6\text{H}_5\text{OH}(\text{s})$?

- (a) Calculate the molar heat of combustion of phenol in kilojoules per mole at 25°C.

$$2.000 \text{g C}_6\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_6\text{H}_5\text{OH}}{94.108 \text{g}} = 0.02125 \text{ mol C}_6\text{H}_5\text{OH} \quad \Delta H_{\text{combustion}} = \frac{-64.98 \text{ kJ}}{0.02125 \text{ mol}} = \boxed{-3058 \frac{\text{kJ}}{\text{mol}}}$$

- (b) Calculate the standard heat of formation, ΔH_f° , of phenol in kilojoules per mole at 25°C.

$$\begin{aligned} \Delta H_{\text{combustion}} &= 6(\Delta H_f^\circ \text{CO}_2) + 3(\Delta H_f^\circ \text{H}_2\text{O}) - \Delta H_f^\circ \text{C}_6\text{H}_5\text{OH} \\ -3058 &= 6(-393.5) + 3(-285.85) - x \\ x &= \boxed{-160.55 \frac{\text{kJ}}{\text{mol}}} \end{aligned}$$