Thermodynamics
6.8 Enthalpy of Formation
Worksheet Key

1) Using standard enthalpy of formation values from the appendix in your textbook, calculate the enthalpy of combustion, ΔH^{o}_{comb}, of one mole of ethane at 25°C.
$C_2H_6(g) + 3.5\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(l)$

$$\Delta H^{o}_{rxn} = \sum \Delta H^{o}_{f\,(products)} - \sum \Delta H^{o}_{f\,(reactants)}$$
$$\Delta H^{o}_{rxn} = [2(\Delta H^{o}_{f}(CO_{2(g)})) + 3(\Delta H^{o}_{f}(H_2O_{(l)}))] - [\Delta H^{o}_{f}(C_2H_{6(g)})]$$
$$\Delta H^{o}_{rxn} = [2(\text{-}393.5\ \text{kJ/mol}) + 3(\text{-}285.83\ \text{kJ/mol})] - (\text{-}84.68\ \text{kJ/mol})$$
$$\Delta H^{o}_{rxn} = -1560\ \text{kJ/mol}$$

2) Using standard enthalpy of formation values from the appendix in your textbook, calculate the enthalpy of reaction, ΔH^{o}_{rxn}, for the formation of carbon dioxide from graphite and oxygen gas. $C_{(graphite)} + O_2(g) \rightarrow CO_2(g)$

$$\Delta H^{o}_{rxn} = \sum \Delta H^{o}_{f\,(products)} - \sum \Delta H^{o}_{f\,(reactants)}$$
$$\Delta H^{o}_{rxn} = [\Delta H^{o}_{f}(CO_{2(g)})] - [(\Delta H^{o}_{f}(C)) + 3(\Delta H^{o}_{f}(O_{2(g)}))]$$
$$\Delta H^{o}_{rxn} = [\Delta H^{o}_{f}(CO_{2(g)})] - 0 = -393.5\ \text{kJ/mol}$$

3) Using standard enthalpy of formation values from the appendix in your textbook, calculate the enthalpy change, ΔH^{o}, for the formation of water vapor from liquid water. $H_2O(l) \rightarrow H_2O\ (g)$

$$\Delta H^{o} = \sum \Delta H^{o}_{f\,(products)} - \sum \Delta H^{o}_{f\,(reactants)}$$
$$\Delta H^{o} = [\Delta H^{o}_{f}(H_2O_{(g)})] - [(\Delta H^{o}_{f}(H_2O_{(l)})]$$
$$\Delta H^{o} = (-241.826\ \text{kJ/mol}) - (-285.840\ \text{kJ/mol})$$
$$\Delta H^{o} = 44.014\ \text{kJ/mol}$$

4) Using standard enthalpy of formation values from the appendix in your textbook, calculate the enthalpy change, ΔH^{o}, for the formation of liquid water from liquid water vapor. $H_2O(g) \rightarrow H_2O\ (l)$

$$\Delta H^{o} = \sum \Delta H^{o}_{f\,(products)} - \sum \Delta H^{o}_{f\,(reactants)}$$
$$\Delta H^{o} = [\Delta H^{o}_{f}(H_2O_{(l)})] - [(\Delta H^{o}_{f}(H_2O_{(g)})]$$
$$\Delta H^{o} = (-285.840\ \text{kJ/mol}) - (-241.826\ \text{kJ/mol})$$
$$\Delta H^{o} = -44.014\ \text{kJ/mol}$$

5) The standard heat of formation of $B_2O_3(s)$ is -1263.6 kJ/mol and the standard heat of formation of liquid water is -285.8 kJ/mol. Use the information below to find the standard heat of formation for one mole of B_5H_9.

$$2\ B_5H_9(l) + 12\ O_2(g) \rightarrow 5\ B_2O_3(s) + 9\ H_2O(l) \qquad \Delta H^\circ_{rxn} = -9036.6\text{ kJ}$$

$$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_{f\,(products)} - \sum \Delta H^\circ_{f\,(reactants)}$$

$$\Delta H^\circ_{rxn} = [5(\Delta H^\circ_f(B_2O_{3(s)})) + 9(\Delta H^\circ_f(H_2O_{(l)}))] - [2(\Delta H^\circ_f(B_5H_{9(g)})) + 12(\Delta H^\circ_f(O_{2(g)}))]$$

$$-9036.6\text{ kJ} = [5(-1263.6\text{ kJ}) + 9(-285.8\text{ kJ})] - [2(\Delta H^\circ_f(B_5H_{9(g)})) + 12(0)]$$

$$-9036.6\text{ kJ} = -8890.2\text{ kJ} - 2(\Delta H^\circ_f(B_5H_{9(g)}))$$

$$\Delta H^\circ_f(B_5H_{9(g)}) = 73.2\text{ kJ/mol}$$

6) Use the following information to find the standard heat of formation for one mole of $NO_2(g)$.

$$2\ NO_2(g) + 7\ H_2(g) \rightarrow 2\ NH_3(g) + 4\ H_2O(l) \qquad \Delta H^\circ = -1303.2\text{ kJ}$$

Standard heat of formation of liquid water $\quad \Delta H^\circ_f = -285.8\text{ kJ / mol}$

Standard heat of formation of gaseous ammonia $\quad \Delta H^\circ_f = -46\text{ kJ / mol}$

$$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_{f\,(products)} - \sum \Delta H^\circ_{f\,(reactants)}$$

$$\Delta H^\circ_{rxn} = [2(\Delta H^\circ_f(NH_{3(g)})) + 4(\Delta H^\circ_f(H_2O_{(l)}))] - [2(\Delta H^\circ_f(NO_{2(g)})) + 7(\Delta H^\circ_f(H_{2(g)}))]$$

$$-1303.2\text{ kJ} = [2(-46\text{ kJ}) + 4(-285.8\text{ kJ})] - [2(\Delta H^\circ_f(NO_{2(g)})) + 7(0)]$$

$$-1303.2\text{ kJ} = -92\text{ kJ} - 1143.2\text{ kJ} - 2(\Delta H^\circ_f(NO_{2(g)}))$$

$$\Delta H^\circ_f(NO_{2(g)}) = 34\text{ kJ/mol}$$

7) Using standard enthalpy of formation values from the appendix in your textbook, calculate the enthalpy of combustion, ΔH°_{comb}, for one mole of methanol at 25°C.

$$CH_3OH(g) + 1.5\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$$

$$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_{f\,(products)} - \sum \Delta H^\circ_{f\,(reactants)}$$

$$\Delta H^\circ_{rxn} = [(\Delta H^\circ_f(CO_{2(g)})) + 2(\Delta H^\circ_f(H_2O_{(l)}))] - [\Delta H^\circ_f(CH_3OH_{(g)}) + 1.5(\Delta H^\circ_f(O_{2(g)}))]$$

$$\Delta H^\circ_{rxn} = [(-393.5\text{ kJ}) + 2(-285.83\text{ kJ})] - [-201.2\text{ kJ} + 1.5(0)]$$

$$\Delta H^\circ_{rxn} = -764.5\text{ kJ/mol}$$

8) Using standard enthalpy of formation values from the appendix in your textbook, calculate the enthalpy of combustion, ΔH°_{comb}, for one mole of C_2H_2 at 25°C.

$$C_2H_2(g) + 1.5\ O_2(g) \rightarrow 2\ CO_2(g) + H_2O(l)$$

$$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_{f\,(products)} - \sum \Delta H^\circ_{f\,(reactants)}$$
$$\Delta H^\circ_{rxn} = [2(\Delta H^\circ_f(CO_{2(g)})) + \Delta H^\circ_f(H_2O_{(l)})] - [\Delta H^\circ_f(C_2H_{2(g)}) + 1.5(\Delta H^\circ_f(O_{2(g)}))]$$
$$\Delta H^\circ_{rxn} = [2(-393.5\text{ kJ}) + (-285.83\text{ kJ})] - [226.7\text{ kJ} + 1.5(0)]$$
$$\Delta H^\circ_{rxn} = -1299.5\text{ kJ/mol}$$

9) Using standard enthalpy of formation values from the appendix in your textbook, calculate the standard enthalpy of combustion, ΔH°_{comb}, of one mole of glucose according to the chemical equation below.

$$C_6H_{12}O_{6(g)} + 6\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 6\ H_2O_{(l)}$$

$$\Delta H^\circ_{rxn} = \Sigma\Delta H^\circ_{f(products)} - \Sigma\Delta H^\circ_{f(reactants)}$$
$$\Delta H^\circ_{rxn} = [6(\Delta H^\circ_f(CO_{2(g)})) + 6(\Delta H^\circ_f(H_2O_{(l)}))] - [\Delta H^\circ_f(C_6H_{12}O_6)]$$
$$\Delta H^\circ_{rxn} = [6\ (-393.5\text{kJ/mol}) + 6\ (-285.83\text{kJ/mol})] - (-1260\text{kJ/mol})$$
$$\Delta H^\circ_{rxn} = -2820\text{ kJ/mol}$$

10) Humans started using elemental copper about 6000 years ago and started using elemental tin about 3800 years ago. Use the heat of formation values for copper (II) oxide and tin (IV) oxide to help explain why humans were able to use elemental copper before they were able to use elemental tin.

Both of these metals exist as oxides in nature, so humans had to transform these oxides into elemental metals.

$$Cu(s) + 1/2\ O_2(g) \rightarrow CuO(s) \qquad H^o_f = -156\text{ kJ/mol}$$

$$Sn(s) + O_2(g) \rightarrow SnO_2(s) \qquad H^o_f = -581\text{ kJ/mol}$$

Since the forward reactions are exothermic, the reverse reactions are endothermic. Thus, both of these metal oxides must be heated in order to drive off the oxygen and produce the elemental metals. It requires + 581 kJ/mol to decompose SnO_2, and only + 156 kJ/mol to decompose CuO. Thus, it requires more energy to decompose SnO_2. As years went by, human cultures got better at concentrating heat from fires, ovens, and furnaces. As these technologies improved, they developed the ability to decompose metal oxides with more negative enthalpy of formation values.

11) Write the balanced chemical equation that outlines the reaction used to determine the enthalpy of formation for one mole of water.

$$H_2(g) + \tfrac{1}{2}\ O_2(g) \rightarrow H_2O(g)$$

12) Write the balanced chemical equation that outlines the reaction used to determine the enthalpy of formation for one mole of $H_2CO(g)$.

$$H_2(g) + \tfrac{1}{2}\ O_2(g) + C(s) \rightarrow H_2CO(g)$$