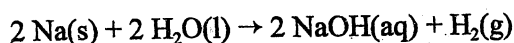
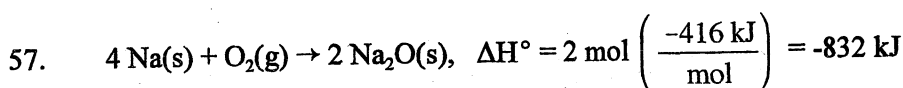


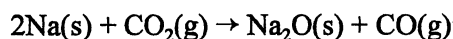
Assignment 4
Week of September 28, 2009
Solutions

Chapter 9: Energy, Enthalpy, and Thermochemistry
57, 62, 64, 86

Standard Enthalpies of Formation

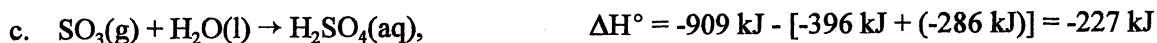
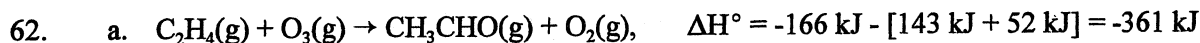


$$\Delta H^\circ = \left[2 \text{ mol} \left(\frac{-470. \text{ kJ}}{\text{mol}} \right) \right] - \left[2 \text{ mol} \left(\frac{-286 \text{ kJ}}{\text{mol}} \right) \right] = -368 \text{ kJ}$$



$$\Delta H^\circ = \left[1 \text{ mol} \left(\frac{-416 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-110.5 \text{ kJ}}{\text{mol}} \right) \right] - \left[1 \text{ mol} \left(\frac{-393.5 \text{ kJ}}{\text{mol}} \right) \right] = -133 \text{ kJ}$$

In reactions 2 and 3, sodium metal reacts with the "extinguishing agent." Both reactions are exothermic and each reaction produces a flammable gas, H₂ and CO, respectively.

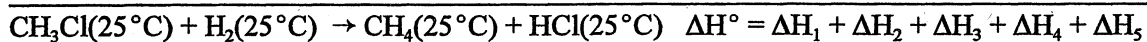
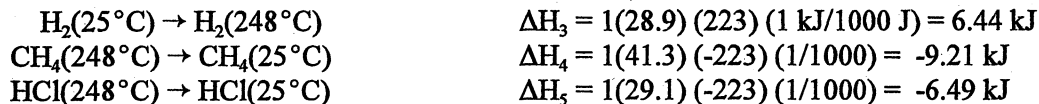
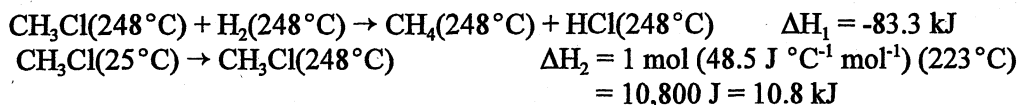


$$\Delta H^\circ = -1411.1 \text{ kJ} = 2(-393.5) \text{ kJ} + 2(-285.8) \text{ kJ} - \Delta H_{\text{f}, \text{C}_2\text{H}_4}^\circ$$

$$-1411.1 \text{ kJ} = -1358.6 \text{ kJ} - \Delta H_{\text{f}, \text{C}_2\text{H}_4}^\circ, \quad \Delta H_{\text{f}, \text{C}_2\text{H}_4}^\circ = 52.5 \text{ kJ/mol}$$

Challenge Problem

86. a. Using Hess's Law and the equation $\Delta H = nC_p\Delta T$:



$$\Delta H^\circ = -83.3 \text{ kJ} + 10.8 \text{ kJ} + 6.44 \text{ kJ} - 9.21 \text{ kJ} - 6.49 \text{ kJ} = -81.8 \text{ kJ}$$

$$\text{b. } \Delta H^\circ = [\Delta H_f^\circ(\text{CH}_4) + \Delta H_f^\circ(\text{HCl})] - [\Delta H_f^\circ(\text{CH}_3\text{Cl}) + \Delta H_f^\circ(\text{H}_2)]$$

$$-81.8 \text{ kJ} = -75 \text{ kJ} - 92 \text{ kJ} - [\Delta H_f^\circ(\text{CH}_3\text{Cl}) + 0], \quad \Delta H_f^\circ(\text{CH}_3\text{Cl}) = -85 \text{ kJ/mol}$$

Chapter 10: Spontaneity, Entropy, and Free Energy
34, 35, 40, 47, 48

Entropy and the Second Law of Thermodynamics: Free Energy

34. Living organisms need an external source of energy to carry out these processes. Green plants use the energy from sunlight to produce glucose from carbon dioxide and water by photosynthesis. In the human body, the energy released from the metabolism of glucose helps drive the synthesis of proteins. For all processes combined, ΔS_{univ} must be greater than zero (2nd law).
35. No, living organisms need an outside source of matter (food) to survive.
40. a. Decrease in disorder; $\Delta S^\circ(-)$ b. Increase in disorder; $\Delta S^\circ(+)$
c. Decrease in disorder ($\Delta n < 0$); $\Delta S^\circ(-)$ d. Decrease in disorder ($\Delta n < 0$); $\Delta S^\circ(-)$
e. HCl(g) is more disordered than two mol of ions in solution; $\Delta S^\circ(-)$.
f. Increase in disorder; $\Delta S^\circ(+)$

For c, d and e, concentrate on the gaseous products and reactants. When there are more gaseous product molecules than gaseous reactant molecules ($\Delta n > 0$), then ΔS° will be positive (disorder increases). When Δn is negative then ΔS° is negative (disorder decreases).

47. a. $\text{NH}_3(\text{s}) \rightarrow \text{NH}_3(\text{l}); \Delta G = \Delta H - T\Delta S = 5650 \text{ J/mol} - 200. \text{ K} (28.9 \text{ J K}^{-1} \text{ mol}^{-1})$

$$\Delta G = 5650 \text{ J/mol} - 5780 \text{ J/mol} = -130 \text{ J/mol}$$

Yes, NH_3 will melt since $\Delta G < 0$ at this temperature.

b. At the melting point, $\Delta G = 0$ so $T = \frac{\Delta H}{\Delta S} = \frac{5650 \text{ J/mol}}{28.9 \text{ J K}^{-1} \text{ mol}^{-1}} = 196 \text{ K}$.

48. a. $S_{\text{rhombic}} \rightarrow S_{\text{monoclinic}}$; This phase transition is spontaneous ($\Delta G < 0$) at temperatures above 95°C . $\Delta G = \Delta H - T\Delta S$; For ΔG to be negative only above a certain temperature, then ΔH is positive and ΔS is positive (see Table 10.6 of text).

b. Since ΔS is positive, then S_{rhombic} is the more ordered crystalline structure.