

Homework Chapter 20

Wednesday, January 10, 2007
12:11 AM

Problems 14, 15, 18, 19, 20, 21, 33, 43, 46, 47, 50, 51, 55, 57, 58, 59, 63, 67, 68, 74, 78, 82, 96, 105

14 a $\Delta S > 0$

b $\Delta S > 0$

c $\Delta S > 0$

15 a $\Delta S < 0$

b $\Delta S < 0$

c $\Delta S < 0$

18 a $\Delta S^\circ > 0$

b $\Delta S^\circ < 0$

c $\Delta S^\circ > 0$

19 a $\Delta S^\circ < 0$

b $\Delta S^\circ > 0$

c $\Delta S^\circ < 0$

20 a $\Delta S^\circ > 0$

b $\Delta S^\circ < 0$

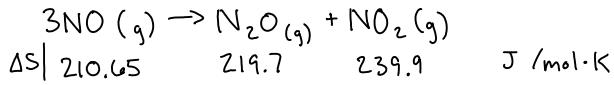
c $\Delta S^\circ > 0$

21 a $\Delta S^\circ < 0$

b $\Delta S^\circ > 0$

c $\Delta S^\circ < 0$

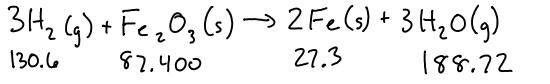
33 a $\Delta S^\circ < 0$



$$219.7 + 239.9 - 3 \times 210.65$$

$$= 172.35 \text{ J/K}$$

b $\Delta S^\circ > 0$

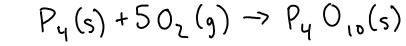


$$(2 \times 27.3 + 3 \times 188.72) - (3 \times 130.6 + 87.400)$$

$$620.76 - 479.2$$

$$= 141.56 \text{ J/K}$$

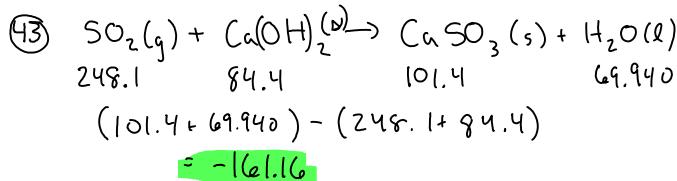
c $\Delta S^\circ < 0$



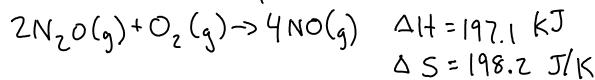
ΔS°	41.1	205.0	229
	J / mol · K		

$$229 - (41.1 + 5 \times 205.0)$$

$$= -837.1 \text{ J/K}$$

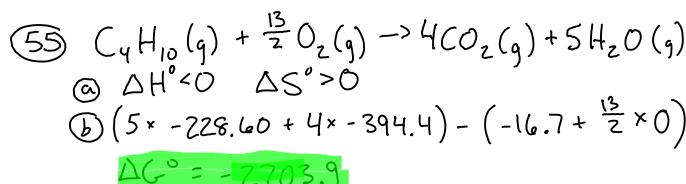
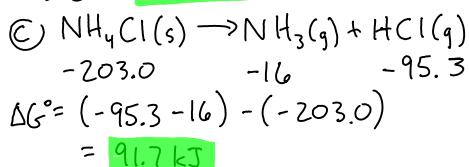
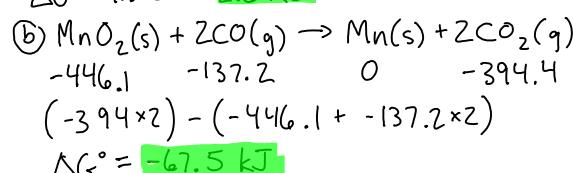
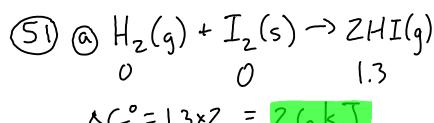
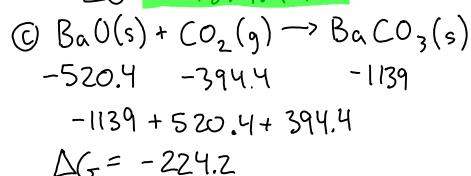
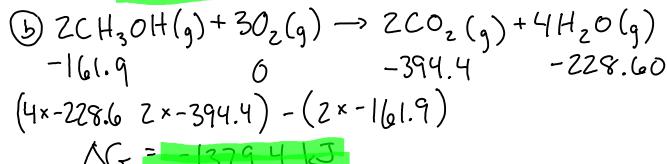
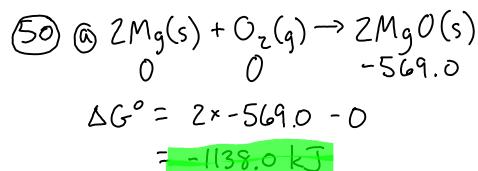


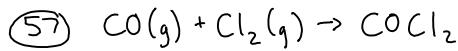
48) ΔH and ΔS are positive.



46) as the spontaneity or entropy increases the free energy decreases

47) higher temperature





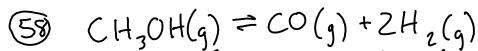
④ $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$$\Delta S^\circ = \frac{\Delta H^\circ - \Delta G^\circ}{T}$$

$$\Delta S^\circ = \frac{(-220. \text{ kJ/mol}) - (-206 \text{ kJ/mol})}{298}$$

$$= -0.0470 \text{ kJ/K}$$

⑤ $\Delta G = -220. \text{ kJ/mol} - (450. \text{ K})(-0.0470 \text{ kJ/K})$
 $= -217.9 \text{ kJ/mol}$



④ $\Delta H^\circ = (2 \times 0 + -110.5) - (-201.2)$
 $= 90.7 \text{ kJ}$

$$\Delta S^\circ = (2 \times 130.6 + 197.5) - (238)$$
 $= 220.7 \text{ J/K}$

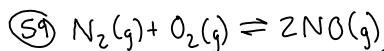
⑤ $\frac{220.7 \text{ J}}{\text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = .2207 \text{ kJ/K}$

$$\Delta G^\circ = 90.7 \text{ kJ} - (311 \text{ K})(.2207 \text{ kJ/K})$$
 $= 22.1 \text{ kJ at } 38^\circ\text{C}$

$$\Delta G^\circ = 90.7 \text{ kJ} - (411 \text{ K})(.2207 \text{ kJ/K})$$
 $= -0.0077 \text{ kJ at } 138^\circ\text{C}$

$$\Delta G^\circ = 90.7 \text{ kJ} - (511 \text{ K})(.2207 \text{ kJ/K})$$
 $= -22.1 \text{ kJ at } 238^\circ\text{C}$

⑥ not spontaneous at 38°C , at equilibrium at 138°C and spontaneous at 238°C



④ $\Delta H^\circ = (2 \times 90.29) - (0 + 0)$
 $= 180.58 \text{ kJ}$

$$\Delta S^\circ = (2 \times 210.65) - (191.5 + 205.0)$$
 $= 24.8 \text{ J/K}$

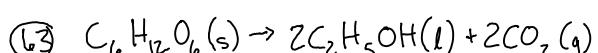
⑤ $\frac{24.8 \text{ J}}{\text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = .0248 \text{ kJ/K}$

$$\Delta G^\circ = 180.58 \text{ kJ} - (373 \text{ K})(.0248 \text{ kJ/K})$$
 $= -171.3 \text{ kJ at } 100^\circ\text{C}$

$$\Delta G^\circ = 180.58 \text{ kJ} - (2833 \text{ K})(.0248 \text{ kJ/K})$$
 $= 110.3 \text{ at } 2560^\circ\text{C}$

$$\Delta G^\circ = 180.58 \text{ kJ} - (3813)(.0248 \text{ kJ/K})$$
 $= 86.0 \text{ kJ at } 3540^\circ\text{C}$

⑦ the reaction is not spontaneous as temperature increases.



$$\Delta H^\circ = (2 \times -393.5 + 2 \times -277.63) - (-1273.3)$$

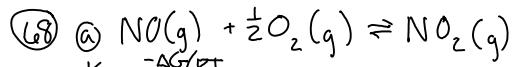
$$\Delta S^\circ = (2 \times 213.7 + 2 \times 161) - (212.1) \\ = \frac{537.3 \text{ J}}{\text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = .5373 \text{ kJ/K}$$

$$\Delta G^\circ = -68.96 \text{ kJ} - (298)(.5373) \\ = -229.1 \text{ kJ}$$

it is spontaneous at all temp. because $\Delta H^\circ < 0$ and $\Delta S > 0$

(67) ΔG° is when all components of the system are in their standard states.

ΔG does not need all components to be in their standard states so temperature, concentration, and pressure can deviate from STP.

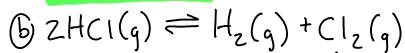


$$K = e^{-\Delta G^\circ / RT}$$

$$\Delta G^\circ = (51) - (86.60 + 0) \\ = -35.6 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = -35600 \text{ J}$$

$$K = e^{-35600/(8.314)(298)}$$

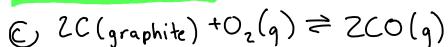
$$K = 1.74 \times 10^4$$



$$\Delta G^\circ = 0 - (-95.30 \times 2) \\ = 190.6 \text{ kJ} = 190600 \text{ J}$$

$$K = e^{-190600/(8.314)(298)}$$

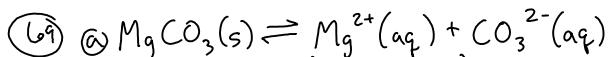
$$K = 3.89 \times 10^{-34}$$



$$\Delta G^\circ = 2 \times -137.2 = -274.4 \text{ kJ} = -274400 \text{ J}$$

$$K = e^{-274400/(8.314)(298)}$$

$$K = 1.26 \times 10^{-49}$$

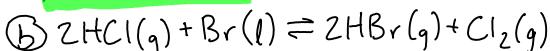


$$\Delta G^\circ = (-528.10) - (-1028)$$

$$= 499.9 \text{ kJ} = 499900 \text{ J}$$

$$K = e^{-499900/(8.314 \times 298)}$$

$$K = 2.36 \times 10^{-89}$$

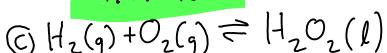


$$\Delta G^\circ = (2 \times -53.5) - (2 \times -95.30)$$

$$= -297.6 \text{ kJ} = -297600 \text{ J}$$

$$K = e^{-297600/(8.314 \times 298)}$$

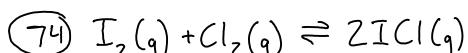
$$K = 1.47 \times 10^{52}$$



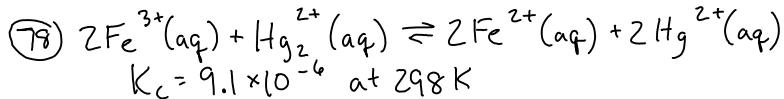
$$\Delta G^\circ = -120.4 \text{ kJ} = -120400 \text{ J}$$

$$K = e^{-120400/(8.314 \times 298)}$$

$$K = 1.27 \times 10^{21}$$



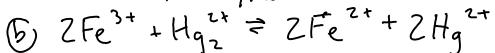
$$\begin{aligned}\Delta G^\circ &= (-6.075 \times 2) - (19.38) \\ &= -31.53 \text{ kJ} = -31530 \text{ J} \\ K &= e^{\frac{-31530}{RT}} \\ &= 3.36 \times 10^5\end{aligned}$$



$$K_c = 9.1 \times 10^{-6} \text{ at } 298 \text{ K}$$

$$\textcircled{a} \Delta G^\circ = -RT \ln K$$

$$= 28758 \text{ J/mol}$$



$$\text{I: } 1\text{M} \quad 1\text{M} \quad 1\text{M} \quad 1\text{M}$$

$$\text{F: } 2-x \quad 1-x \quad 2+x \quad 2+x$$

$$9.1 \times 10^{-6} = \frac{(2+x)^2(2+x)^2}{(2-x)^2(1-x)}$$

$$= \frac{(4+2x+x^2)(4+2x+x^2)}{(4-2x+x^2)(1-x)} = \frac{16 + 8x + 4x^2 + 8x + 4x^2 + 2x^3 + 4x^2 + 2x^3 + x^4}{4 - 2x + x^2 - 4x + 2x^2 - x^3}$$

$$= \frac{x^4 + 4x^3 + 12x^2 + 16x + 16}{-x^3 + 3x^2 - 6x + 4}$$

or neglect some x s

$$\frac{16 + 8x + 4x^2}{4} = 9.1 \times 10^{-6}$$

$$16 + 8x + 4x^2 = 3.64 \times 10^{-5}$$

$$4x^2 + 8x + 15.9999636$$

$$\frac{-8 \pm \sqrt{64 - 255.99}}{8} \text{ negative } \sqrt$$



$$\textcircled{a} \Delta H^\circ = -1.896 \text{ kJ/mol}$$

$$\Delta G^\circ = -2.866 \text{ kJ/mol}$$

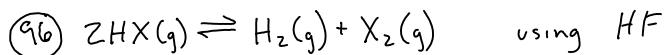
$$\Delta S^\circ = 5.686 - 2.439 = 3.247 \text{ J/mol}\cdot\text{K}$$

$$\textcircled{b} K = e^{\frac{-2866}{RT}} \\ = 3.456$$

diamonds are not forever, perhaps many lifetimes, but not forever.

C 1.896 kJ/mol would be required at 298 K

D it is not spontaneous for the reverse reaction $\Delta H > 0 \Delta S < 0$



$$\textcircled{a} \Delta H^\circ = 2 \times -273 = -546 \text{ kJ/mol}$$

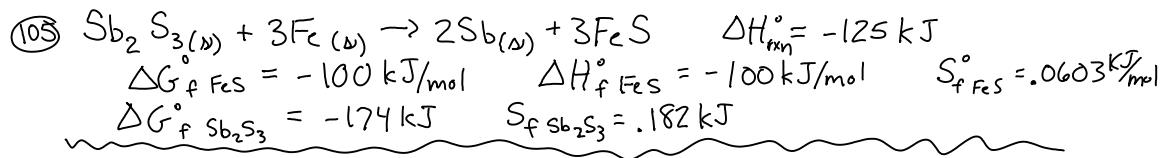
$$\Delta S^\circ = (130.6 + 202.7) - (2 \times 173.67) = -14.04 \text{ J/mol}\cdot\text{K}$$

$$\Delta G^\circ = 2 \times -273 = -550 \text{ kJ/mol}$$

B reaction is spontaneous at low temp because $\Delta H^\circ < 0 \Delta S^\circ < 0$

C no

D when $\Delta G = 0$



② $\Delta H_{rxn}^\circ = \sum \Delta H_f^{\circ} \text{ prod} - \sum \Delta H_f^{\circ} \text{ reactants}$
 $-125 = [2(\Delta H_f^{\circ} Sb_2S_3) + 3(\Delta H_f^{\circ} FeS)] - [\Delta H_f^{\circ} Sb_2S_3 + 3\Delta H_f^{\circ} Fe]$
 $-125 = 0 + 3(-100 \text{ kJ/mol}) - x - 3(0)$ $x = \Delta H_f^{\circ} Sb_2S_3$
 $x = 425 \text{ kJ} = \Delta H_f^{\circ} Sb_2S_3$

③ $\Delta G_{rxn}^\circ = (3x - 100 + 2 \cdot 0) - (3 \cdot 0 - 174)$
 $= -126 \text{ kJ/mol}$

④ S° for $Sb(s)$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$-126 \text{ kJ/mol} = -125 \text{ kJ/mol} - (298)(\Delta S)$$

$$-1 \text{ kJ/mol} = -298 \Delta S$$

$$\Delta S^\circ = 3.356 \times 10^{-3}$$

$$3.356 \times 10^{-3} = (.0603 \times 3 + 2x) - (.0273 \times 3 + .182)$$

$$3.356 \times 10^{-3} = 2x - .083$$

$$x = -.0398 \text{ kJ} = \Delta S_{Sb(s)}$$