

Chapter 6 Homework Questions

- 6.1** The standard Gibbs Energy of the reaction, $\text{H}_2(\text{g}) + 3 \text{N}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$, is -32.9 kJ/mol at 25°C . Calculate the value of ΔG for the following values of the reaction quotient, Q : (a) 0.01, (b) 1.0, (c) 10.0, (d) 100,000, (e) 1,000,000
Estimate the value of the equilibrium constant, K , from the above data, and calculate the actual value of K .
- 6.2** Consider the gas phase reaction, $2 \text{H}_2\text{O}(\text{g}) \rightleftharpoons 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g})$. At 2257 K and $p = 2.0 \text{ bar}$, H_2O is 1.77% dissociated. Calculate the equilibrium constant for this reaction.
- 6.3** Consider the reaction, $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$. The Gibbs Energy of Formation of $\text{NH}_3(\text{g})$ is -16.5 kJ/mol at 298 K .
Calculate the reaction Gibbs Energy for the above reaction (at 298 K) when the partial pressures of the gases are; $p_{\text{N}_2} = 3.0 \text{ bar}$, $p_{\text{H}_2} = 1.0 \text{ bar}$, $p_{\text{NH}_3} = 0.50 \text{ bar}$. What is the spontaneous direction of the reaction under these conditions?
- 6.4** Consider the gas phase equilibrium, $3 \text{A}(\text{g}) \rightleftharpoons \text{B}(\text{g}) + 2 \text{C}(\text{g})$. At 50°C , the Gibbs Energy change for the reaction is -2.50 kJ/mol when the component pressures are: $p_{\text{A}} = 2.0 \text{ bar}$, $p_{\text{B}} = 4.0 \text{ bar}$, $p_{\text{C}} = 0.50 \text{ bar}$,
Calculate the value of the equilibrium constant, K , at 50°C .
- 6.5** Consider the gas phase dissociation equilibrium, $\text{A}(\text{g}) \rightleftharpoons \text{B}(\text{g}) + 3 \text{C}(\text{g})$. At 25°C , the percent dissociation is 30% at a total pressure of 2 bar. Calculate the equilibrium constant, K .
- 6.6** Consider the dissociation equilibrium, $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g})$. At 25°C and $p = 1 \text{ bar}$, N_2O_4 is 18.46% dissociated. The reaction enthalpy is constant at $\Delta H^\circ = +56.2 \text{ kJ/mol}$.
(a) Calculate the equilibrium constant, K , at 25°C
(b) Calculate K at 100°C
- 6.7** For the gas phase reaction, $2 \text{A} + \text{B} \rightleftharpoons 3 \text{C} + 2 \text{D}$, it was found that when 1.00 mol of A, 2.00 mol of B, and 1.00 mol of D were mixed and allowed to come to equilibrium at 25°C , the resulting mixture contained 0.90 mol of C at a total pressure of 4.0 bar.
(a) Calculate the equilibrium constant, K , for this reaction
(b) Calculate the Reaction Gibbs Energy, $\Delta_r G^\circ$, for this reaction at 25°C .
- 6.8** The standard reaction enthalpy for the reaction, $\text{Zn}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{ZnO}(\text{s}) + \text{H}_2(\text{g})$, is approximately constant at $224. \text{ kJ/mol}$. The standard reaction Gibbs energy is $+33 \text{ kJ/mol}$ at 1280 K .
Determine the temperature at which the Equilibrium Constant for this reaction becomes greater than 1.

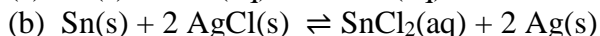
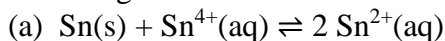
- 6.9** The equilibrium constant for the reaction, $A(g) \rightleftharpoons B(g)$, is temperature dependent and follows the equation:

$$\ln K = A + \frac{B}{T} + \frac{C}{T^2}, \quad A = -1.0, \quad B = -1100 \text{ K}, \quad C = 1.5 \times 10^5 \text{ K}^2$$

Calculate

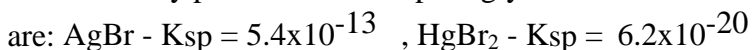
- (a) the Reaction Enthalpy, ΔH° (in kJ/mol) at 180 °C
(b) the Reaction Entropy, ΔS° (in J/mol K) at 180 °C
- 6.10** For the gas phase equilibrium, $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$, the equilibrium constant is $K = 4.5$ at 150 °C.
- (a) If one puts $PCl_5(g)$ into a container with an initial pressure of $P_0 = 3.0$ bar, what are the pressures of $PCl_5(g)$ and $PCl_3(g)$ at equilibrium.
- (b) Determine the Gibbs energy change for the when the partial pressures of the three gases are: $P(PCl_5) = 0.8$ bar and $P(PCl_3) = 2.2$ bar and $P(Cl_2) = 2.2$ bar?
- 6.11** For the gas phase equilibrium, $2A(g) + B(g) \rightleftharpoons 2C(g)$, the equilibrium constant is $K = 1 \times 10^{-3}$. If one puts A and B into a vessel with initial pressures, $P_{init}(A) = P_{init}(B) = 3.$ bar, determine the pressure of C at equilibrium?
NOTE: You may assume that very little A and B react.
- 6.12** Consider the equilibrium, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. The value of the equilibrium constant is $K = 337$ at 100 °C and $K = 7.1 \times 10^{-5}$ at 500 °C
- (a) Determine the enthalpy (ΔH°) and entropy (ΔS°) changes for this reaction.
- (b) Determine the value of the equilibrium constant at 300 °C.
- (c) At 500 °C, the initial pressures (before reaction to form NH_3) of N_2 and H_2 are 3.0 bar and 2.0 bar, respectively. Calculate the pressure of NH_3 at equilibrium [Note: You can make the assumption that very little NH_3 is formed relative to the initial pressures of N_2 and H_2 to simplify your calculation].
- (d) Determine the Gibbs energy change for the reaction at 500 °C for $P_{N_2} = P_{H_2} = 0.2$ bar and $P_{NH_3} = 2.5$ bar

6.13 Use the Reduction Potential Table below to calculate the equilibrium constants for the following two reactions:



Reduction Reaction	E_{red}°
$\text{Sn}^{4+}(\text{aq}) + 2 e^{-} \rightarrow \text{Sn}^{2+}(\text{aq})$	+0.15 V
$\text{Sn}^{2+}(\text{aq}) + 2 e^{-} \rightarrow \text{Sn(s)}$	-0.14 V
$2 \text{AgCl(s)} + 2 e^{-} \rightarrow 2 \text{Ag(s)} + 2 \text{Cl}^{-}(\text{aq})$	+0.22

6.14 The solubility products of two sparingly soluble Bromide (Br^{-}) salts



Consider a solution which initially contains $5.0 \times 10^{-5} \text{ M Ag}^{+}(\text{aq})$ and $5.0 \times 10^{-5} \text{ M Hg}^{2+}(\text{aq})$.

KBr (a strong electrolyte) is added until $[\text{Br}^{-}] = 2.0 \times 10^{-8} \text{ M}$. Which of the above salts (AgBr and HgBr_2) will precipitate?

6.15 Mercury(I) Sulfate, Hg_2SO_4 , is a sparingly soluble salt with $K_{\text{sp}} = 6.5 \times 10^{-7}$.

If 1200 mL of 0.010 M $\text{K}_2\text{SO}_4(\text{aq})$ is mixed with 800 mL of 0.020 M $\text{HgNO}_3(\text{aq})$, calculate the concentrations of $[\text{Hg}^{+}]$ and $[\text{SO}_4^{2-}]$ in the resulting solution and **determine whether or not** $\text{Hg}_2\text{SO}_4(\text{s})$ will precipitate.

6.16 The standard reduction potentials of $\text{Br}_2(\text{l})$ and $\text{Hg}^{2+}(\text{aq})$ are +1.07 V and +0.86 Volts, respectively Consider the electrochemical cell,



(a) Write the balanced equation for this reaction, and indicate the number of electrons transferred.

(b) Calculate the standard cell potential, E°_{cell} , for the reaction.

(c) Calculate the equilibrium constant for the reaction at 25 °C.

(d) Calculate the cell potential under the conditions shown above.

6.17 The standard reduction potentials of $\text{Mg}^{2+}(\text{aq})$ and $\text{Al}^{3+}(\text{aq})$ are -2.37 V and -1.66 Volts, respectively. Consider the electrochemical cell,



(a) Write the balanced equation for this reaction, and indicate the number of electrons transferred.

(b) Calculate the standard cell potential, E°_{cell} , for the reaction.

(c) Calculate the equilibrium constant for the reaction at 25 °C.

(d) Calculate the cell potential under the conditions shown above.

- 6.18** The EPA recommended maximum concentration of Zn^{2+} [$M(\text{Zn}) = 65.4 \text{ g/mol}$] in drinking water is 5. mg/L. The amount of Zn in a sample of water can be determined by measuring the voltage of an electrochemical cell in which the reference electrode (cathode) has a standard concentration [say, 0.20 M $\text{Zn}(\text{NO}_3)_2$] and the sample electrode (anode) has the water sample. This cell can be designated as: $\text{Zn(s)}|\text{Zn}^{2+}(\text{xx M})||\text{Zn}^{2+}(0.20 \text{ M})|\text{Zn(s)}$.

The cell potential was measured as +0.078 V. Determine the concentration of Zn^{2+} in the sample, in mg/L.

- 6.19** An electrochemical cell is prepared with 0.50 M $\text{Pb}(\text{NO}_3)_2(\text{aq})$ in the reference compartment (cathode) and a saturated solution of lead iodate, $\text{Pb}(\text{IO}_3)_2(\text{aq})$, in the sample compartment (anode). The measured cell voltage is: 0.120 V.

Calculate the Solubility Product, K_{sp} , of $\text{Pb}(\text{IO}_3)_2$.

- 6.20** An electrochemical cell is prepared with 0.25 M $\text{AgNO}_3(\text{aq})$ in the reference compartment (cathode) and a saturated solution of silver phosphate, $\text{Ag}_3\text{PO}_4(\text{aq})$, in the sample compartment (anode). The measured cell voltage is: 0.195 V.

Calculate the Solubility Product, K_{sp} , of Ag_3PO_4 .

- 6.21** Consider the electrochemical cell reaction: $\text{Hg}_2\text{Cl}_2(\text{aq}) + \text{H}_2(\text{g}) \rightleftharpoons 2 \text{Hg}(\text{l}) + 2 \text{HCl}(\text{aq})$. The standard cell potential is temperature dependent and given by:

$E^\circ_{\text{cell}} = a - b/T^2$ where $a = 0.313 \text{ V}$ and $b = 5.0 \times 10^7 \text{ V/K}^2$. , T is temperature in Kelvins.

For this reaction, calculate: (a) $\Delta_r G^\circ$, (b) $\Delta_r S^\circ$, (c) $\Delta_r H^\circ$