Equilibrium HW Holt May 2017 **Answer Key**

p. 595 (PP 1-3, SR 1-10), p. 604 (SR 1-6);
p. 616 (PP 1&2); p. 618 (PP 1&2); p. 620 (PP 1&2, SR 1-7)
pp. 622-624 (2-11, 14-16, 27, 29, 32, 33, 34, 37, 39, 40 (*review in class problems*)

P 595 (PP 1-3)

1. At equilibrium a mixture of N_2, H_2 and NH_3 gas at 500°C is determined to consist of 0.602 mol/L of N_2, 0.420 mol/L of H_2, and 0.113 mol/L of NH_3. What is the equilibrium constant for the reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ at this temperature?

$$K_{eq} = [NH_3]^2/[N_2]\ [H_2]^3 = (0.113)^2/\ (0.602)\ (0.420)^3 = \mathbf{0.286}$$

2. The reaction $AB_2C\ (g) \rightarrow B_2\ (g) + AC\ (g)$ reached equilibrium at 900 K in a 5.00 vessel. At equilibrium 0.084 mol of AB_2C, 0.035 mol of B2, and 0.059 mol of AC were detected What is the equilibrium constant at this temperature for this system? (Don't forget to convert amounts to concentrations.) $K_{eq} = [B_2]\ [AC]/\ [AB_2C] = (0.007)(0.0118)\ /\ 0.0168) = \mathbf{4.9 \times 10^{-3}}$

3. A reaction between gaseous sulfur dioxide and oxygen gas to produce gaseous sulfur trioxide takes place at 600°C. At that temperature, the concentration of SO_2 is found to be 1.50 mol/L, the concentration of O_2 is 1.25 mol/L, and the concentration of SO_3 is 3.50 mol/L. Using the balanced chemical equation, calculate the equilibrium constant for this system. $2SO_2 + O_2 \Leftrightarrow 2SO_3$

$K_{eq} = [SO3]^2\ /\ [SO_2]^2\ [O_2] = (3.50)2\ /\ (1.50)2\ (1.25) = \mathbf{4.36}$

P. 595 (SR 1-10)

1. What is meant by chemical equilibrium? The rates of forward and reverse chemical reactions are equal, and the concentrations of products and reactants remain unchanged.
2. What is an equilibrium constant? The ration of the mathematical product of concentrations of substances formed at equilibrium to the mathematical product of the reactant concentrations, each raised to a power equal to the corresponding coefficient in the balanced chemical equation.
3. How does the value of an equilibrium constant relate to the relative quantities of reactants and products at equilibrium? The larger the value of *K* is, the larger the relative amount of products.
4. What is meant by a chemical equilibrium expression?
   The expression of product and reactant concentrations that equal an equilibrium constant, *K*
5. Hydrochloric acid, HCl, is a **strong acid that dissociates completely** in water to form H_3O^+ and Cl^-. Would you expect the value of K for the reaction $HCl\ (aq) + H_2O\ (l) \rightarrow H_3O^+(aq) + Cl^-(aq)$ to be 1×10^{-2}, 1×10^{-5}, or "very large"? Justify your answer. Very large, because virtually all of the reactants have formed products
6. Write the chemical equilibrium expression for the reaction $4HCl\ (g) + O_2\ (g) \rightarrow 2Cl_s(g) + 2H_2O(g)$

   $$K_{eq} = [Cl_2]^2[H_2O]^2{}_/\ [HCl]^4[O_2]$$

7. At equilibrium at 2500K, [HCl] = 0.0625 mol/L and $[H_2] = [Cl_2] = 0.00450$ mol/L for the reaction $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$. Find the value of K.

$$K_{eq} = [HCl]^2/\ [H_2]\ [Cl_2] = (0.0625)^2/(0.0045)^2 = \mathbf{193}$$

8. An equilibrium mixture at 425°C is found to consist of 1.83×10^{-3} mol/L of H_2, 3.13×10^{-3} mol/L of I_2, and 1.77×10^{-2} mol/L of HI. Calculate the equilibrium constant, K, for the reaction $H_2(g) + I_2(g) \rightarrow 2HI(g)$.

$$K_{eq} = [HI]^2/\ [H_2]\ [I_2]= (1.77 \times 10^{-2})^2\ /\ (1.83 \times 10^{-3})(3.13 \times 10^{-3}) = \mathbf{54.7}$$

9. For the reaction H_2 (g) + I_2(g) → 2HI(g) at 425°C, calculate [HI], given $[H_2] = [I_2] = 4.79 \times 10^{-4}$ mol/L and K = 54.3.

$K_{eq} = [HI]^2 / [H_2]\,[I_2]$ so $[HI]^2 = K_{eq}[H_2][I_2]$ and then square root =

*$54.3\,(4.79 \times 10^{-4})^2$ = **3.53×10^{-3} mol/L***

10. Using the data from experiment 1 in Table 1 to calculate the value of K for the reaction 2HI (g) → H_2 g) + I_2(g). Do you see a relationship between the value you obtained and the value in the table?

*$Keq = [H_2]\,[I_2] / [HI]^2 = (0.4953 \times 10^{-3})^2/(3.655 \times 10^{-3})^2$ = **0.184** which is the inverse of 54.46 because it is the reverse reaction.*

p. 604 (SR1-6)

1. Name three ways the chemical equilibrium can be disturbed. *Change in concentration, change in pressure (volume), and change in temperature.*

2. Describe three situations in which ionic reaction go to completion. *Formation of gas, a ppt, or a slightly ionized product.*

3. Describe the common-ion effect. *The addition of an ion common to two solutes brings about precipitation or reduced ionization.*

4. Identify the common ion in each of the following situations.

   a. 5 g of NaCl is added to a 2.0 M solution of HCl. **Cl^-**

   b. 50 ml of 1.0 M $NaCH_3COO$ is added to 1.0 M CH_3COOH. **CH_3COO^-**

   c. 10 pellets of NaOH are added to 100 ml of water. **OH^-**

5. Predict the effect that decreasing pressure would have on each of the following reaction systems at equilibrium.

   a. H_2 (g) + Cl_2 (g) ⇔ 2HCl (g) ***No effect***

   b. NH_4Cl (s) ⇔ NH_3 (g) + HCl (g) ***shift right***

   c. $2H_2O_2$ (aq) ⇔ $2H_2O$ (l) + O_2 (g) ***shift right***

   d. $3O_2$ (g)⇔ $2O_3$ (g) ***shift left***

6. Carbon dioxide and water react to form bicarbonate ion and hydronium ion. Hyperventilation (rapid breathing) causes more carbon dioxide toe be exhaled than normal. How will hyperventilation affect the pH of the blood? Explain.

CO_2 (g) + H_2O (l) ⇔ HCO_3^- (aq) + H_3O^+ (aq). As CO_2 is lost, this equilibrium system shifts left, decreasing hydronium concentration, thus causing the pH to increase.

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p. 616 (PP 1 & 2)

1. Calculate the solubility product constant, K_{sp}, of lead (II) chloride, $PbCl_2$, which has a solubility of 1.0 g/100.g H_2O at 20°C.

   $PbCl_2$ (s) ⇔ Pb^{2+} (aq) + $2Cl^-$ (aq) $K_{sp} = [Pb^{2+}][Cl^-]^2$

   1.0g/ 277.9 g ÷ 0.100 l = 0.036 M (0.036)(0.072)2 = $\underline{1.86 \times 10^{-4}}$

2. A 5.0 gram sample of Ag_2SO_4 will dissolve in 1.0 L of water. Calculate the solubility product constant for this salt. ***5.0 g Ag_2SO_4/311.74 g/mol ÷ 1.0 liter = 0.016 M***

Ag_2SO_4 (s) ⇔ $2Ag^+$ (aq) +SO_4^{2-}(aq) $K_{sp} = [Ag^+]^2\,[SO_4^{2-}]$ = (0.032)2 (0.016) = $\underline{1.6 \times 10^{-5}}$

p. 618 (PP 1 & 2)

1. Calculate the solubility of cadmium sulfide, CdS, in mol/L given the K_{sp} value listed in Table 3.

   $[Cd^{2+}][S^{2-}]$ = 8.0 x 10-27 $[Cd^{2+}] = [S^{2-}]$ = ***$\underline{8.9 \times 10^{-14}\text{ mol/L}}$***

2. Determine the concentration of strontium ions in a saturated solution of strontium sulfate, $SrSO_4$ if the K_{sp} for $SrSO_4$ is 3.2 x 10^{-7}.

*$SrSO_4$ (s) ⇔ Sr^{2+}(aq)+SO_4^{2-} (aq);K_{sp} = [Sr^{2+}][SO_4^{2-}] = 3.2 x 10^{-7} so [Sr^{2+}]= **5.65 x 10^{-4} M***

p. 620 (PP 1 & 2)

1. Does precipitate from when 100. mL of 0.0025 M $AgNO_3$ and 150. mL of 0.0020 M NaBr solutions are mixed?

 $AgNO_3$ M_1V_1/V_t = (0.0025M)(0.1 L)/(0.25 L) = 0.001 M Ag^+ and NO_3^-
 NaBr M_1V_1/V_t = (0.0020M)(0.150 L)/(0.25 L) = 0.0012 M Na^+ and Br^-
 AgBr ⇔ Ag+ (aq) + Br- (aq) Ksp = [Ag^+][Br^-] = (0.001)(0.0012) = 1.2 x 10^{-6} = Q_{sp}
 Q_{sp} > K_{sp} (5.0 x 10^{-13}) so **AgBr precipitates**

2. Does a precipitate form when 20. mL of 0.0038 M $Pb(NO_3)_2$ and 30. mL of 0.018 M KCl solutions are mixed? $PbCl_2$ does not precipitate.

 $Pb(NO_3)_2$ M_1V_1/V_t = (0.0038)(0.020 l)/(0.050 l) = 0.00152 M $Pb(NO_3)_2$ → Pb^{2+} + 2NO_3^-
 KCl M_1V_1/V_t = (0.018)(0.030)/(0.050) = 0.0108 M KCl → K^+ + C^{l-}
 Qsp = [Pb^{2+}][Cl^-]^2 = (0.00152)(0.0108)^2 = 1.77 x 10^{-7} **Q_{sp} < K_{sp}** (1.6 x 10^{-5}) so **$PbCl_2$ does not precipitate**

p. 620 (SR 1-7)

1. What is a solubility product constant? How are such constants determined? The product of the molar concentrations of the ions in a saturated solution, each raised to a power that is the coefficient of that ion in the chemical equation; from careful measurements of ion solubilities.
2. How are solubility product constants used to calculate solubilities? By using the solubility equilibrium equation, solving for an ion concentration, and then relating ion concentration to moles of solute.
3. What is an ion product? The product of the ion concentrations present in solution, each raised to the appropriate power.
4. How are calculations to predict possible precipitation carried out? The ion product is calculated; if it exceeds the K_{sp} then a precipitate will form.
5. What is the value of K_{sp} for Ag_2SO_4 if 5.40 g is soluble in 1.00 L of water? *5.4 g Ag_2SO_4/311.74 g/mol ÷ 1.0 liter = 0.017 M K_{sp} = [Ag^+]^2 [SO_4^{2-}] = (0.034)^2((0.017) = **1.96 x 10^{-5} M*** *book:* 2.07 x 10^{-5}
6. Determine whether a precipitate will form if 20.0 mL of 1.00 x 10^{-7} M $AgNO_3$ is mixed with 20.0 mL of 2.00 x 10^{-9} M NaCl at 425°C.

 $AgNO_3$ (aq) + NaCl (aq) → AgCl (s) + 2$NaNO_3$ (aq)
 AgCl (s) ⇔ Ag^+ (aq) + Cl^- (aq) Ksp = [Ag^+][Cl^-] = 1.8 x 10^{-10}
 (1.0x 10^{-7} M)(0.02 L) = M_2 (0.040 L) → M_2 = 5.0 x 10^{-8} M [Ag^+]
 (2.0 x 10^{-9} M)(0.02 L) = M_2 (0.040 L) → M_2 = 1.0 x 10-9 M [Cl^-]
 Q = (5.0 x 10^{-8} M) (1.0 x 10^{-9} M) = **5.0 x 10^{-17} < 1.8 x 10^{-10} so NO PPT will form**

7. A solution is 0.20 M in each of the following: $Ca(NO_3)_2$, $Cr(NO_3)_3$, and $La(NO_3)_3$. Solid NaF is added to the solution until the [F^-] of the solution is 1.0 x 10^{-4} M. Given the values of K_{sp}, describe what will happen. CaF_2 = 3.9 x 10^{-11}; CrF_3 = 6.6 x 10^{-11}; and LaF_3 = 4.0 x 10^{-17}

 CaF_2 (s) ⇔ Ca2+ (aq) + 2F^- (aq) Q_{sp} = [Ca^{2+}][F^-]^2 (0.20 M)(1 x 10^{-4} M)^2 = 2.0 x 10^{-9} > K_{sp} **ppt**
 CrF_3 (s) ⇔ Cr^{3+} (aq) + 3F^- (aq) Q_{sp} = [Cr^{3+}][F^-]^3 (0.20 M)(1 x 10^{-4} M)^3 = 2.0 x 10^{-13} < K_{sp} no ppt
 LaF_3 (s) ⇔ La^{3+} (aq) + 3F^- (aq) Q_{sp} = [La^{3+}][F^-]^3 (0.20 M)(1 x 10^{-4} M)^3 = 2.0 x 10^{-13} > K_{sp} **ppt**

p. 622-624 (2-11, 14-16, 27, 29, 32, 33, 34, 37, 39,40)

2. a. Write the general expression for an equilibrium constant based on the equation

$nA + mB \Leftrightarrow xC + yD$ $\quad$ $K_{eq} = \frac{[C]^x[D]^y}{[A]^n[B]^m}$

b. What information is provided by the value of K for a given equilibrium systems at a specified temperature? ***It tells us the relative amounts of products vs. reactants***

3. In general, which reaction is favored (forward, reverse, or neither) if the value of *K* at a specified temperature is
   a. Very small? ***Reverse***
   b. Very large? ***Forward***
4. Determine the value of the equilibrium constant for each reaction given, assuming that the equilibrium concentrations are found to be those specified. (concentrations are given in mol/L) (Hint: see Sample Problem A)

a. $A + B \Leftrightarrow C$; $\quad$ [A] = 2.0; [b] = 3.0; [C] = 4.0 $\quad$ $K_{eq} = \frac{[C]}{[A]\,[B]} = 4.0/(2.0)(3.0) = \underline{\mathbf{0.67}}$

b. $D + 2E \Leftrightarrow F + 3G$;[D] = 1.5; [E]=2.0;[F]=1.8;[G]=1.2 $\quad$ $K_{eq} = \frac{[F][G]^3}{[D\,[E]^2} = \frac{(1.8)(1.2)^3}{(1.5)(2.0)^2} = \underline{\mathbf{0.52}}$

c. $N_2(g)+3H_2(g) \Leftrightarrow 2NH_3(g)$;$[N_2]$=0.45;$[H_2]$=0.14;$[NH_3]$=0.62

$$K_{eq} = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} = \frac{(0.62)^2}{(0.45)(0.14)^3} = \underline{\mathbf{3.1 \times 10^2}}$$

5. An equilibrium mixture at a specific temperature is found to consist of $1.2\text{x}10^{-3}$ mol/L HCI, $3.8\text{x}10^{-4}$ mol/L O_2, $5.8\text{x}10^{-2}$mol/L H_2O, and $5.8\text{x}10^{-2}$ mol/L Cl_2 according to the following: $4HCI(g)+ O_2(g) \Leftrightarrow 2H_2O(g)+ 2CI_2(g)$. Determine the value of the equilibrium constant for this system.

$$K_{eq} = \frac{[H_2O]^2\,[Cl_2]^2}{[HCl]^4\,[O_2]} = \frac{(5.8\text{x}10^{-2}M)^2(5.8\text{x}10^{-2}M)^2}{(1.2\text{x}10^{-3}M)^4(\,3.8\text{x}10^{-4}M)} = \underline{\mathbf{1.4 \times 10^{10}}}$$

6. At 450 ºC, the value of the equilibrium constant for the following system is $6.59\text{x}\ 10^{-3}$. If $[NH_3]$ = 1.23×10^{-4} M and $[H_2]$= $2.75\text{x}\ 10^{-2}$M at equilibrium, determine the concentration of N_2 at that point. $N_2(g)+ 3H_2(g) \Leftrightarrow 2NH_3(g)$

$$K_{eq} = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} \qquad [N_2] = \frac{[NH_3]^2}{K_{eq}[H_2]^3} = \frac{(1.23 \times 10^{-4}\ M)^2}{(6.59 \times 10^{-3})(2.75 \times 10^{-2}\ M)^3} = \underline{\mathbf{0.110\ M}}$$

7. The value of the equilibrium constant for the reaction below is 40.0 at a specified temperature. What would be the value of that constant for the reverse reaction under the same conditions?
   $H_2(g)+ I_2(g) \Leftrightarrow 2HI(g)$; *it would be the reciprocal of 40.0 =* ***0.025***

8. Predict whether each of the following pressure changes would favor the forward or reverse reaction. $2NO(g)+ O_2(g) \Leftrightarrow 2NO_2(g)$
   a. Increased pressure ***to the right***
   b. Decreased pressure ***to the left***

9. In heterogeneous reaction systems, what types of substances do not appear in the equilibrium constant expression? Why? ***Liquids and solids as their concentrations are their densities which do not vary with temperature***

10. Explain the effect of a catalyst on an equilibrium system. ***A catalyst has NO effect on an equilibrium system as it increases the rate of both forward and reverse reactions, but has no effect on equilibrium concentrations of reactants and products***

11. Predict the effect of each of the following on the indicated equilibrium system in terms of the direction of equilibrium shift (forward, reverse, or neither)

H_2 (g) + Cl_2 (g) ⇔ 2HCl (g) + 184 kJ

a. Addition of Cl_2 ***shift right***

b. Removal of HCl ***shift right***

c. Increased pressure ***no effect; moles of reactants = moles of products***

d. Decreased temperature ***shift right***

e. Removal of H_2 ***shift left***

f. Decreased pressure ***no effect; moles of reactants = moles of products***

g. Addition of a catalyst ***No effect***

h. Increased temperature ***shift left***

i. Decreased system volume ***no effect; moles of reactants = moles of products***

14. What relative pressure (high or low) would result in the production of the maximum level of CO_2 according the following equation? Why? 2CO (g) + O_2 (g) ⇔ 2CO_2 (g) ***high pressure because the forward reaction converts three gas molecules into two, relieving the stress imposed by the pressure increase***

15. What relative conditions (reactant concentrations, pressure, and temperature) would favor a high equilibrium concentration of the underlined substance in each of the following equilibrium systems?

a. 2CO (g) + O_2 (g) ⇔ $\underline{2CO_2\ (g)}$ + 167 kJ ***high reactant [ ], high pressure, low temperature***

b. Cu^{2+} (aq) + 4NH_3 (aq) ⇔ $\underline{Cu(NH_3)_4^{2+}}$ + 42 kJ ***high reactant [ ], pressure NA, low temperature***

c. 2HI (g) + 112.6 kJ ⇔ H_2 (g) + $\underline{I_2\ (g)}$ ***high reactant [ ], pressure NA, high temperature***

d. 4HCl(g) + O_2 (g) ⇔ 2H_2O (g) + $\underline{2Cl_2\ (g)}$ + 113 kJ ***high reactant [ ], high pressure, low temperature***

e. PCl_5 (g) + 88 kJ ⇔ PCl_3 (g) + $\underline{Cl_2(g)}$ ***high reactant [ ], low pressure, high temperature***

16. The reaction between hemoglobin, Hb, and oxygen, O_2, in red blood cells is responsible for transporting O_2 to body tissues. This process can be represented by the following equilibrium reaction: Hb (aq) + O_2(g) ⇔ HbO_2 (aq) What will happen to the concentration of oxygenated hemoglobin, HbO_2, at high altitude, where the pressure of oxygen is 0.1 atm instead of 0.2 atm, as it is at sea level? ***The concentration of oxygenated hemoglobin will be less than it is at sea level.***

27. What is the relationship between Ksp and the product of the ion concentrations in terms of determining whether a solution of those ions is saturated? ***Q < Ksp, unsaturated and ppt will not occur; if Q> Ksp, ppt will occur until ion concentrations decrease to equilibrium values.***

29. Calculate the solubility product constant, Ksp for each of the following, based on the solubility information provided:

a. $BaSO_4$ = 2.4 x 10^{-4} g/100 g. H_2O at 20 °C. *(text answer - **1.1 x 10^{-10}**)*

*$BaSO_4$ (s) ⇔ Ba^{2+} (aq) + SO_4^{2-} (aq); Ksp = [Ba^{2+}][SO_4^{2-}]= (1.03 x 10^{-5} M)^2 =**1.06 x 10^{-10}***

2.4 x 10^{-4} g/100 g H_2O x 1 g/ml H_2O x 1000 ml/1 L H_2O x 1 mol $BaSO_4$/ 233 g $BaSO_4$ = 1.03 x 10^{-5} M

b. $Ca(OH)_2$ = 0.173 g/100 g. H_2O at 20 °C.

*$Ca(OH)_2$ (s)⇔Ca^{2+}(aq) +2OH^- (aq); Ksp = [Ca^{2+}][OH^-]^2 = (2.34 x 10^{-2} M)(**2** x 2.34 x 10^{-2} M)^2= **5.12 x 10^{-5}***

0.173 g/100 g H_2O x 1 g/ml H_2O x 1000 ml/1 L H_2O x 1 mol $Ca(OH)_2$/ 74 g $Ca(OH)_2$ = 2.34 x 10^{-2} M

32. Complete each of the following relative to the reaction that occurs when 25.0 ml of 0.0500 M $Pb(NO_3)_2$ is combined with 25.0 ml of 0.0400 M Na_2SO_4 if equilibrium is reached at 25 °C.

a. write the solubility equilibrium equation at 25 °C.

$Pb(NO_3)_2$ (aq) + Na_2SO_4 (aq) → $NaNO_3$ (aq) + $PbSO_4$ (s)

$PbSO_4$ (s) ⇔ Pb^{2+} (aq) + SO_4^{2-} (aq)

b. write the solubility equilibrium expression for the net reaction.

K_{sp} = [Pb^{2+}][SO_4^{2-}]

33. The ionic substance T_3U_2 ionizes to form T^{2+} and U^{3-} ions. The solubility of T_3U_2 is 3.8 x 10^{-10} mol/L. What is the value of the solubility product constant?

T_3U_2 (s) ⇔ 3T^{2+}(aq) + 2U^{3-} (aq); Ksp = [T^{2+}]^3 [U^{3-}]^2=(3 x 3.8 x 10^{-10})^3 (2 x 3.8 x 10^{-10})^2 = 8.56 x 10^{-46}

34. A solution of AgI contains 2.7 x 10^{-10} mol/L Ag^+. What is the maximum I^- concentration that can exist in this solution? ***[K_{sp} = 8.3 x 10^{-17} (p. 615)]***

*AgI (s) ⇔ Ag^+ (aq) + I^- (aq); K_{sp} = [Ag^+][I^-] = **8.3 x 10^{-17} = (2.7 x 10^{-10})(x) → x = 3.07 x 10^{-7} M***

37. If 2.50 x 10^{-2} g of solid $Fe(NO_3)_3$ is added to 100. ml of a 1.0 x 10^{-4} M NaOH solution, will a precipitate form? ***K_{sp} = 4 x 10^{-38}***

$Fe(OH)_3$ (s) ⇔ Fe^{3+} (aq) + 3OH^- (aq)

[Fe^{3+}] = 2.5 x 10^{-2} g/242 g/mol ÷ 0.10 L = 1.03 x 10^{-3} M

K_{sp} = [Fe^{3+}][OH^-]^3 = (1.03 x 10^{-3} M) (1.0 x 10^{-4} M)^3 = 1.03 x 10^{-15} (Q)> 4 x 10^{-38}

Q > Ksp so a ppt will form

39. Calculate the concentrations of Hg^{2+} ions in a saturated solution of HgS(s). How many Hg^{2+} ions are in 1000 L of the solution?

HgS (s) ⇔ Hg^{2+} (aq) + S^{2-} (aq); K_{sp} = [Hg^{2+}][S^{2-}] = 1.6 x 10^{-52}; [Hg^{2+}] = 1.26 x 10^{-26} M

1.26 x 10^{-26} M x 1000 L x 6.02 x 10^{23} ions/mol = 7.6 ~ 8 ions

40. Calculate the equilibrium constant, *K*, for the following reaction at 900 ºC.

H_2(g) + CO_2(g) ⇔ H_2O (g) + CO (g) The components were analyzed, and it was found that [H_2] = 0.061 M, [CO_2] = 0.16 M, [H_2O] =0.11 M, and [CO] = 0.14 M.

$$K_{eq} = \frac{[H_2O]\,[CO]}{[H_2]\,[CO_2]} = \frac{(0.11)\,(0.14)}{(0.061)\,(0.16)} = 1.6$$