REVIEW QUESTIONS
Chapter 14

1. A mixture of 0.10 mol of NO, 0.050 mol of H_2 and 0.10 mol of H_2O is placed in a 1.0-L flask and allowed to reach equilibrium as shown below:

$$2\ NO\ (g) + 2\ H_2\ (g) \rightleftharpoons N_2\ (g) + 2\ H_2O\ (g)$$

At equilibrium [NO] = 0.062 M. Calculate the equilibrium constant, K_c, for this reaction.

	2 NO +	2 H_2 (g) $\rightleftharpoons$	N_2 (g) +	2 H_2O (g)
Initial	**0.10 M**	**0.050 M**	**0**	**0.10 M**
Δ	**– 0.038**	**– 0.038**	**+ 0.019**	**+ 0.038**
Equilibrium	**0.062**	**0.012**	**0.019**	**0.138**

$$\mathbf{K_C = \frac{[N_2][H_2O]^2}{[NO]^2[H_2]^2} = \frac{(0.019)(0.138)^2}{(0.062)^2(0.012)^2} = 650}$$

2. At 700°C, K_c = 20.4 for the reaction shown below:

(A) $SO_2\ (g) + ½\ O_2\ (g) \rightleftharpoons SO_3\ (g)$

Calculate K_c and K_P for the reaction shown below:

(B) $2\ SO_2\ (g) + O_2\ (g) \rightleftharpoons 2\ SO_3\ (g)$

$$\mathbf{K_C\ (B) = [K_C\ (A)]^2 = (20.4)^2 = 416}$$

$$\mathbf{K_P = K_C\ (RT)^{\Delta n} = (416)\ [(0.0821)(973\ K)]^{-1} = 5.21}$$

3. At 100°C, K_c= 0.078 for the following reaction:

$$SO_2Cl_2\ (g) \rightleftharpoons SO_2\ (g) + Cl_2\ (g)$$

In an equilibrium mixture, $[SO_2Cl_2]$ = 0.136 M and $[SO_2]$ = 0.072 M. What is the concentration of Cl_2 in the equilibrium mixture?

$$\mathbf{K_C = \frac{[SO_2][Cl_2]}{[SO_2Cl_2]} = 0.078}$$

$$\mathbf{[Cl_2] = \frac{0.078\ [SO_2Cl_2]}{[SO_2]} = \frac{(0.078)(0.136)}{0.072} = 0.15\ M}$$

4. At 373 K, K_P= 0.416 for the equilibrium:

$$2\ NOBr\ (g) \rightleftharpoons 2\ NO\ (g) + Br_2\ (g)$$

If the partial pressures of NOBr and NO are equal at equilibrium, what is the partial pressure of Br_2?

$$K_p = \frac{\cancel{P^2_{NO}}\ P_{Br_2}}{\cancel{P^2_{NOBr}}} = 0.416 \qquad \text{since } P_{NOBr} = P_{NO}$$

$$P_{Br_2} = 0.416\ \text{atm}$$

5. At 250°C, the reaction

$$PCl_5\ (g) \rightleftharpoons PCl_3\ (g) + Cl_2\ (g)$$

has an equilibrium constant K_c= 1.80. If 0.100 mol of PCl_5 is added to a 5.00-L flask, what are the concentrations of PCl_5, PCl_3 and Cl_2 at equilibrium at this temperature?

$$[PCl_5] = \frac{0.100\ \text{mol}}{5.00\ \text{L}} = 0.0200\ \text{M}$$

	$PCl_5\ (g) \rightleftharpoons$	$PCl_3\ (g)$ +	$Cl_2\ (g)$
Initial	**0.0200 M**	**0**	**0**
Δ	**– x**	**+ x**	**+ x**
Equilibrium	**0.0200 –x**	**x**	**x**

$$K_C = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{x^2}{0.0200 - x} = 1.80$$

$$x^2 + 1.80\,x - 0.0360 = 0$$

Solving the quadratic equation,

$$x = \frac{-1.80 \pm \sqrt{(1.80)^2 - 4(-0.0360)}}{2} = 0.0198$$

$$[PCl_3]=[Cl_2]= 0.0198\ \text{M}$$

$$[PCl_5]= 0.0200 - 0.0198 = 0.0002\ \text{M}$$

6. When 2.00 mol each of hydrogen and iodine are mixed in a 1.00-L flask, 3.50 mol of HI is produced at equilibrium:

$$H_2\,(g) + I_2\,(g) \rightleftharpoons 2\,HI\,(g)$$

Calculate the equilibrium constant K_c for this reaction.

	$H_2\,(g)$ +	$I_2\,(g)$ $\rightleftharpoons$	$2\,HI\,(g)$
Initial	**2.00 M**	**2.00 M**	**0**
Δ	**– 1.75**	**– 1.75**	**+ 3.50**
Equilibrium	**0.25**	**0.25**	**3.50**

$$K_C = \frac{[HI]^2}{[H_2]\,[I_2]} = \frac{(3.50)^2}{(0.25)^2} = 196$$

7. The equilibrium constant for the reaction

$$SO_2\,(g) + NO_2\,(g) \rightleftharpoons SO_3\,(g) + NO\,(g)$$

has a numerical value of 3.00 at a given temperature. 1.50 mol each of SO_2 and NO_2 are mixed in a 1.00-L flask and allowed to reach equilibrium. What percent of SO_2 is converted to product?

	$SO_2\,(g)$ +	$NO_2\,(g)$ $\rightleftharpoons$	$SO_3\,(g)$ +	$NO\,(g)$
Initial	**1.50 M**	**1.50 M**	**0**	**0**
Δ	**– x**	**– x**	**+ x**	**+ x**
Equilibrium	**1.50 –x**	**1.50 –x**	**x**	**x**

$$K_C = \frac{[SO_3][NO]}{[SO_2][NO_2]} = \frac{x^2}{(1.50 - x)^2} = 3.00$$

Taking square root of each side,

$$\frac{x}{1.50 - x} = 1.73$$

$$x + 1.73\,x = 2.5\underline{9}5$$

$$x = 0.951$$

$$\% = \frac{0.951}{1.50} x100 = 63.4\%$$

8. The following equilibrium exists at 1000 ºC with $K_C = 2.00$.

$$2\,COF_2\,(g) \rightleftharpoons CO_2\,(g) + CF_4\,(g)$$

If a 5.00-L mixture contains 0.145 mol COF_2 , 0.262 mol of CO_2 and 0.074 mol of CF_4 at 1000 ºC, in which direction will the mixture proceed to reach equilibrium?

$$[COF_2] = \frac{0.145\text{ mol}}{5.00\text{ L}} = 0.0290\text{ M} \qquad [CO_2] = \frac{0.262\text{ mol}}{5.00\text{ L}} = 0.0524\text{ M}$$

$$[CF_4] = \frac{0.074\text{ mol}}{5.00\text{ L}} = 0.0148\text{ M}$$

$$Q_c = \frac{[CO_2][CF_4]}{[COF_2]^2} = \frac{(0.0524)(0.0148)}{(0.0290)^2} = 0.922$$

Since $Q_c < K_c$, reaction will proceed in the forward direction

9. Formamide decomposes at high temperatures according to the equation shown below:

$$HCONH_2\,(g) \rightleftharpoons NH_3\,(g) + CO\,(g) \qquad K_C = 4.84 \text{ at } 400\text{ K}$$

If 0.186 mol of formamide is placed in a 2.16-L flask and allowed to decompose at 400 K, what will be the total pressure at equilibrium?

$$P_{HCONH_2} = \frac{nRT}{V} = \frac{(0.186)(0.0821)(400\text{ K})}{2.16\text{ L}} = 2.83\text{ atm}$$

$$K_p = K_c\,(RT)^{\Delta n} = 4.84\,[(0.0821)(400)]^1 = 159$$

	$HCONH_2\,(g)$ $\rightleftharpoons$	$NH_3\,(g)$ +	$CO\,(g)$
Initial	**2.83 atm**	**0**	**0**
Δ	**– x**	**+ x**	**+ x**
Equilibrium	**2.83 – x**	**x**	**x**

$$K_p = \frac{P_{NH_3}\,P_{CO}}{P_{HCONH_2}} = \frac{(x)^2}{(2.83 - x)} = 159$$

Solving the quadratic equation, x = 2.78

$$P_{HCONH_2} = 2.83 - 2.78 = 0.05\text{ atm}$$

$$P_{NH_3} = P_{CO} = 2.78\text{ atm}$$

$$P_{Total} = 2.78 + 2.78 + 0.05 = 5.61\text{ atm}$$

10. Predict how each of the following changes affect the amount of H_2 present in an equilibrium mixture in the reaction

$$3\,Fe\,(s) + 4\,H_2O\,(g) \rightleftharpoons Fe_3O_4\,(s) + 4\,H_2\,(g) \qquad \Delta H = -150\text{ kJ}$$

a) Raising the temperature of the mixture.
Since reaction is exothermic, raising temperature will shift the equilibrium to the left (←) and reduce amount of hydrogen.

b) Adding more H_2O (g).
Adding more water, will shift the equilibrium to the right (→) and increase the amount of hydrogen.

c) Doubling the volume of the container holding the mixture.
Increasing the volume of the container will reduce the pressure but the equilibrium will not be affected, and the amount of hydrogen will not change.

d) Adding a catalyst.
Adding catalyst does not alter the equilibrium and the amount of hydrogen produced.

11. At 2000 °C the equilibrium constant for the reaction below is $K_c = 2.4\times10^3$. If the initial concentration of NO is 0.500 M, what are the equilibrium concentrations of each substance?

$$2\,NO\,(g) \rightleftharpoons N_2\,(g) + O_2\,(g)$$

	2 NO (g) ⇌	N_2 (g) +	O_2 (g)
Initial	**0.500 M**	**0**	**0**
Δ	**– 2x**	**+ x**	**+ x**
Equilibrium	**0.500 –x**	**x**	**x**

$$K_C = \frac{[N_2][O_2]}{[NO]^2} = \frac{x^2}{(0.500 - 2x)^2} = 2.4\times10^3$$

Taking square root of each side,

$$\frac{x}{0.500 - 2x} = 49 \qquad [N_2] = [O_2] = x = 0.25\text{ M}$$

$$x + 98\,x = 2\underline{4}.5 \qquad [NO] = 0.500 - 2x = 0.500 - 0.4\underline{9}4$$

$$x = 0.2\underline{4}7 \qquad [NO] = 6.0\times10^{-3}\text{ M}$$

12. The reaction below has an equilibrium constant $K_c = 6.90$. If 0.100 mol of BrCl is placed in a 500-mL flask and allowed to come to equilibrium, what are the equilibrium concentrations of each substance?

$$Br_2\,(g) + Cl_2\,(g) \rightleftharpoons 2\,BrCl\,(g)$$

$$[BrCl] = \frac{0.100\text{ mol}}{0.500\text{ L}} = 0.200\text{ M}$$

	Br_2 (g)	+ Cl_2 (g)	$\rightleftharpoons$ 2 BrCl (g)
Initial	**0**	**0**	**0.200 M**
Δ	**+x**	**+ x**	**–2x**
Equilibrium	**x**	**x**	**0.200 –2x**

$$K_C = \frac{[BrCl]^2}{[Br_2][Cl_2]} = \frac{(0.200 - 2x)^2}{(x)^2} = 6.90$$

Taking square root of each side,

$$\frac{0.200 - 2x}{x} = 2.63$$

$2.63\,x + 2\,x = 0.200$

$x = 0.0432$

$[B_2] = [C_2] = x = 0.0432\text{ M}$

$[BrCl] = 0.200 - 2x = 0.200 - 0.0864$

$[BrCl] = 0.114\text{ M}$