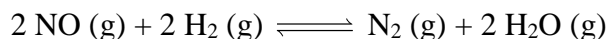## REVIEW QUESTIONS
Chapter 15

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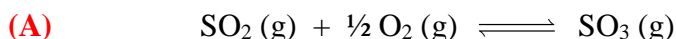
$$2\ NO\ (g) + 2\ H_2\ (g) \rightleftharpoons N_2\ (g) + 2\ H_2O\ (g)$$

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**

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

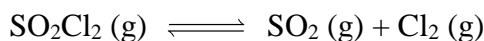2. At 700°C, K_c = 20.4 for the reaction shown below:

   (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)

   K_C (B) = [K_C (A)]² = (20.4)² = 416

   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:

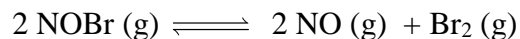$$SO_2Cl_2\ (g) \rightleftharpoons SO_2\ (g) + Cl_2\ (g)$$

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?

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

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

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

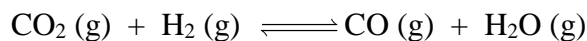$$2 \text{ NOBr (g)} \rightleftharpoons 2 \text{ NO (g)} + \text{Br}_2 \text{ (g)}$$

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

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

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

5. A mixture of 0.100 mol CO_2, 0.0500 mol H_2 and 0.100 mol H_2O are place in a 1.00-L flask and allowed to come to equilibrium as shown below. At equilibrium $[CO_2]$ = 0.0954 M.

$$\text{CO}_2 \text{ (g)} + \text{H}_2 \text{ (g)} \rightleftharpoons \text{CO (g)} + \text{H}_2\text{O (g)}$$

a) What are the equilibrium concentrations of H_2, CO and H_2O?

	CO_2 (g) +	H_2 (g) $\rightleftharpoons$	CO (g) +	H_2O (g)
Initial	0.100	0.0500	0.00	0.100
Δ	− 0.0046	− 0.0046	+ 0.0046	+ 0.0046
Equilibrium	0.0954	0.0454	0.0046	0.1046

b) Calculate K_c and K_p for this reaction at 25.0°C.

$$K_C = \frac{[CO][H_2O]}{[CO_2][H_2]} = \frac{(0.0046)(0.1046)}{(0.0954)(0.0454)} = 0.11$$

$$K_p = K_c(RT)^{\Delta n} = 0.11 \, [(0.0821)(298)]^0 = 0.11$$

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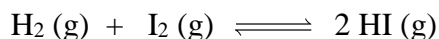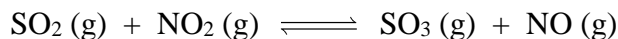
$$H_2 (g) + I_2 (g) \rightleftharpoons 2\ HI (g)$$

Calculate the equilibrium constant K_c for this reaction.

	H₂ (g) +	I₂ (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?

	SO₂ (g) + NO₂ (g) $\rightleftharpoons$ SO₃ (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**

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

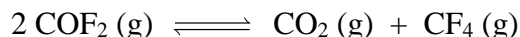**Taking square root of each side,**

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

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

$$x = 0.951$$

$$\% = \frac{0.951}{1.50} \times 100 = 63.4\%$$

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?

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

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

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

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

9. A 0.831-g sample of SO_3 is placed in a 1.00-L flask and heated to 1100 K. The SO_3 decomposes to SO_2 and O_2, as shown below. At equilibrium, the total pressure in the container is 1.300 atm. Find the values of K_p and K_c for this reaction at 1100 K.

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

$$mol\ SO_3 = 0.831\ g\ x\ \frac{1\ mol}{80.06\ g} = 0.01038\ mol$$

$$P_{SO_3} = \frac{nRT}{V} = \frac{(0.01038)(0.0821)(1100\ K)}{1.00\ L} = 0.9374\ atm$$

	2 SO$_3$ (g) $\rightleftharpoons$	2 SO$_2$ (g) +	O$_2$ (g)
Initial	0.9374	0.00	0.00
Δ	− 2x	+ 2x	+ x
Equilibrium	0.9374 − 2x	2x	x

$P_{tot} = P_{SO3} + P_{SO2} + P_{O2} = 1.300\ atm$

$P_{tot} = (0.9374\text{-}2x) + 2x + x = 1.300\ atm$

$x = P_{O2} = 0.3626\ atm \qquad P_{SO2} = 2\ P_{O2} = 0.7252\ atm \qquad P = 0.9374 - 2x = 0.2122\ atm$

$$K_p = \frac{P_{SO2}^2\ P_{O2}}{P_{SO3}^2} = \frac{(0.7252)^2(0.3626)}{(0.2122)^2} = 4.2$$

$$K_c = \frac{K_p}{(RT)^{\Delta n}} = \frac{4.2}{[(0.0821)(1100)]^1} = 0.047$$

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

$$3 \text{ Fe (s)} + 4 \text{ H}_2\text{O (g)} \rightleftharpoons \text{Fe}_3\text{O}_4 \text{ (s)} + 4 \text{ H}_2 \text{ (g)} \qquad \Delta H = -150 \text{ kJ}$$
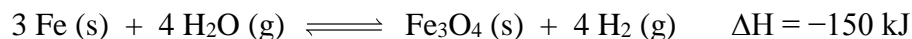
a) Raising the temperature of the mixture.
Since reaction is exothermic, raising temperature will shift the equilibrium to the left ($\leftarrow$) and reduce amount of hydrogen.

b) Adding more H_2O (g).
Adding more water, will shift the equilibrium to the right ($\rightarrow$) 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 $K_c = 2.4 \times 10^3$. If the initial concentration of NO is 0.500 M, what are the equilibrium concentrations of each substance?
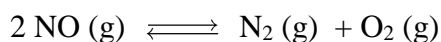
$$2 \text{ NO (g)} \rightleftharpoons \text{N}_2 \text{ (g)} + \text{O}_2 \text{ (g)}$$

	2 NO (g) $\rightleftharpoons$ N₂ (g)	+	O₂ (g)
Initial	**0.500 M**	**0**	**0**
Δ	**– 2x**	**+ x**	**+ x**
Equilibrium	**0.500 –x**	**x**	**x**

$$K_C = \frac{[N_2][O_2]}{[NO]^2} = \frac{x^2}{(0.500 - 2x)^2} = 2.4 \times 10^3$$

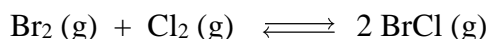**Taking square root of each side,**

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

$$x + 98 \text{ x} = 24.5 \qquad\qquad [NO] = 0.500 - 2x = 0.500 - 0.494$$

$$x = 0.247 \qquad\qquad [NO] = 6.0 \times 10^{-3} \text{ M}$$

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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] = 0.100 mol/0.500 L= 0.200 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,

$$\frac{0.200 - 2x}{x} = 2.63 \qquad [Br_2] = [Cl_2] = x = 0.0432 \text{ M}$$

2.63 x + 2 x = 0.200   **[BrCl] = 0.200 - 2x = 0.200 - 0.0864**

x = 0.0432   **[BrCl] = 0.114 M**

13. An equilibrium mixture of H_2, I_2, and HI at 458°C contains 2.24×10^{-2} M H_2, 2.24×10^{-2} M I_2 and 0.155 M HI in a 5.00-L flask. What are the equilibrium concentrations when equilibrium is reestablished following the addition of 0.100 mol of HI?

$$\text{H}_2 \text{ (g)} + \text{I}_2 \text{ (g)} \rightleftharpoons 2 \text{ HI (g)} \qquad K_C = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.155)^2}{(2.24 \times 10^{-2})^2} = 47.9$$

After addition of 0.100 mol HI: **[HI] = 0.155 M + 0.0200 M) = 0.175 M**

	H_2 (g)	+	I_2 (g) $\rightleftharpoons$	2 HI (g)
Initial	**2.24×10^{-2}**		**2.24×10^{-2}**	**0.175**
Δ	**+x**		**+x**	**−2x**
Equilibrium	**(2.24×10^{-2})+x**		**(2.24×10^{-2})+x**	**0.175–2x**

$$K_C = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.175 - 2x)^2}{[(2.24 \times 10^{-2})+x]^2} = 47.9$$

Taking square root of each side,

$$\frac{0.175 - 2x}{(2.24 \times 10^{-2})+x} = 6.92 \qquad \textbf{0.175 - 2 x = 0.155 + 6.92x}$$

8.92x = 0.020   **x = 0.0022**

$[H_2] = [I_2] = 2.46 \times 10^{-2}$ M   **[HI] = 0.171 M**