CHEM 1423
Chapters 17
Homework Solutions

TEXTBOOK HOMEWORK

17.29 2 HBr(g) ⇌ H_2(g)+Br_2(g)
P_{HBr} = 0.2 atm , P_{H2} = 0.01 atm , P_{Br2} = 0.01 atm

$$Q = \frac{P_{H_2} \cdot P_{Br_2}}{P_{HBr}^2} = \frac{(0.01)(0.01)}{(0.2)^2} = 2.5x10^{-3} > K_P\left(4.18x10^{-9}\right)$$

Because Q ≠ K_P, the reaction is **not** at equilibrium.
Because Q > K_P, the reaction will move towards the left until Q = K_P

17.38 2 NO(g) + Cl_2(g) ⇌ 2 NOCl(g) $K_P = 6.5x10^4$, P_{NO} = 0.35 atm , P_{Cl2} = 0.10 atm

$$K_P = 6.5x10^4 = \frac{P_{NOCl}^2}{P_{NO}^2 \cdot P_{Cl_2}} = \frac{P_{NOCl}^2}{(0.35)^2(0.10)} = \frac{P_{NOCl}^2}{1.225x10^{-2}}$$

$$P_{NOCl}^2 = \left(6.5x10^4\right)(1.225x10^{-2}) = 796.3$$

$$P_{NOCl} = \sqrt{796.3} = 28.2\,atm$$

17.41 2 H_2S(g) ⇌ 2 H_2(g) + S_2(g) , $K_c = 9.3x10^{-8}$, $[H_2S]_o$ = 0.45 mol/3 L = 0.15 M

	H_2S	H_2	S_2
Initial	0.15	0	0
Change	-2x	+2x	+x
Equilibrium	0.15-2x	2x	x

As stated in the problem (see outline), one may assume that very little H_2S will decompose (because K_c is very small). Therefore, at equilibrium,
$[H_2S]$ = 0.15 - 2x ≈ 0.15 M

$$K_c = 9.3x10^{-8} = \frac{[H_2]^2[S_2]}{[H_2S]^2} = \frac{(2x)^2(x)}{(0.15)^2} = 177.8x^3$$

$$x^3 = \frac{9.3x10^{-8}}{177.8} = 5.23x10^{-10}$$

$$x = \left(5.23x10^{-10}\right)^{1/3} = 8.06x10^{-4}\,M$$

$[H_2]$ = 2x = $1.61x10^{-3}$ M $[S_2]$ = x = $8.06x10^{-4}$ M

17.44 2 ICl(g) ⇌ I_2(g) + Cl_2(g) , K_c = 0.110 , $[ICl]_o$ = 0.50 mol/5.0 L = 0.10 M

	ICl	I_2	Cl_2
Initial	0.10	0	0
Change	-2x	+x	+x
Equilibrium	0.10-2x	x	x

Note: Because it is not stated in the problem, you may NOT assume that very little H_2S will decompose.

$$K_c = 0.110 = \frac{[I_2][Cl_2]}{[ICl]^2} = \frac{(x)(x)}{(0.10-2x)^2} = \frac{x^2}{(0.10-2x)^2}$$

Take square root of both sides

$$\sqrt{0.110} = 0.332 = \frac{x}{0.10-2x}$$

$$0.332(0.10-2x) = x$$

$$0.0332 - 0.664x = x$$

$$1.664x = 0.0332$$

$$x = 0.01995 = 0.020$$

$[I_2]$ = $[Cl_2]$ = x = 0.020 M [ICl= 0.10 - 2x = 0.10 - 2(0.02) = 0.060 M

17.46 4 NH_3(g) + 3 O_2(g) ⇌ 2 N_2(g) + 6 H_2O(g)
$[NH_3]_o$ = $[O_2]_o$ = 0.015 mol/1.00 L = 0.015 M , $[N_2]_{equil}$ = 1.96x10^{-3} M

Let's make an ICE Table

	NH_3	O_2	N_2	H_2O
Initial	0.015	0.015	0	0
Change	-4x	-3x	+2x	+6x
Equilibrium	0.015-4x	0.015-3x	2x	6x

We can determine the value of x by using the equilibrium concentration, $[N_2]$.
$[N_2]_{equil}$ = 2x = 1.96x10^{-3} ⇒ x = 9.80x10^{-4}
$[H_2O]$ = 6x = 5.88x10^{-3} , $[NH_3]$ = 0.015-4x = 1.108x10^{-2} , $[O_2]$ = 0.015-3x=1.206x10^{-2}

$$K_c = \frac{[N_2]^2[H_2O]^6}{[NH_3]^4[O_2]^3} = \frac{(1.96x10^{-3})^2\left(5.88x10^{-3}\right)^6}{(1.108x10^{-2})^4(1.206x10^{-2})^3} = 6.01x10^{-6} \approx 6.0x10^{-6}$$

17.47 FeO(s) + CO(g) ⇌ Fe(s) + CO_2(g) K_P = 0.403
Note: We can ignore FeO(s) and Fe(s). Only gases need be considered.

	FeO(s)	CO(g)	Fe(s)	CO_2(g)
Initial	--	1.00 atm	0	0
Change	--	-x	--	+x
Equilibrium	--	1.00-x	--	x

$$K_P = 0.403 = \frac{P_{CO_2}}{P_{CO}} = \frac{x}{1.00 - x}$$

$$0.403(1.00 - x) = x$$

$$0.403 - 0.403x = x$$

$$1.403x = 0.403$$

$$x = \frac{0.403}{1.403} = 0.287\,atm$$

P_{CO} = 1.00-x =1.00-0.287 = 0.713 atm , P_{CO2} = x = 0.287 atm

17.56 An increase in volume results in a decrease in pressure. Therefore, the equilibrium will move in the direction which increases the number of moles of gas. However, there is no change in the equilibrium constant.

(a) F_2(g) ⇌ 2 F(g): Equil. will move to the right. More F and less F_2. No change in K

(b) 2 CH_4(g) ⇌ C_2H_2(g) + 3 H_2(g): Equil. will move to the right. More C_2H_2 and H_2 and less CH_4. No change in K

17.61 When the temperature is **decreased**, the equilibrium will move in the exothermic direction. K will change accordingly.

(a) Exothermic reaction (ΔH^o_{rxn} = -151 kJ). Equilibrium will move to right and K will increase.

(b) Exothermic reaction (ΔH^o_{rxn} = -451 kJ). Equilibrium will move to right and K will increase.

(c) Exothermic reaction. Equilibrium will move to right and K will increase.

(d) Endothermic reaction. Equilibrium will move to left and K will decrease.