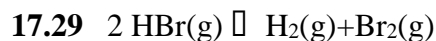


CHEM 1423
Chapters 17
Homework Solutions

TEXTBOOK HOMEWORK

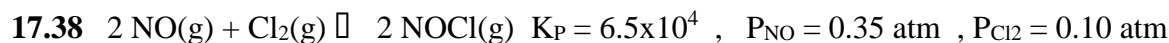


$P_{\text{HBr}} = 0.2 \text{ atm}$, $P_{\text{H}_2} = 0.01 \text{ atm}$, $P_{\text{Br}_2} = 0.01 \text{ atm}$

$$Q = \frac{P_{\text{H}_2} \cdot P_{\text{Br}_2}}{P_{\text{HBr}}^2} = \frac{(0.01)(0.01)}{(0.2)^2} = 2.5 \times 10^{-3} > K_p (4.18 \times 10^{-9})$$

Because $Q \neq K_p$, the reaction is **not** at equilibrium.

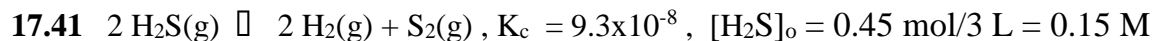
Because $Q > K_p$, the reaction will move towards the left until $Q = K_p$



$$K_p = 6.5 \times 10^4 = \frac{P_{\text{NOCl}}^2}{P_{\text{NO}}^2 \cdot P_{\text{Cl}_2}} = \frac{P_{\text{NOCl}}^2}{(0.35)^2 (0.10)} = \frac{P_{\text{NOCl}}^2}{1.225 \times 10^{-2}}$$

$$P_{\text{NOCl}}^2 = (6.5 \times 10^4)(1.225 \times 10^{-2}) = 796.3$$

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



	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,
 $[\text{H}_2\text{S}] = 0.15 - 2x \approx 0.15 \text{ M}$

$$K_c = 9.3 \times 10^{-8} = \frac{[\text{H}_2]^2 [\text{S}_2]}{[\text{H}_2\text{S}]^2} = \frac{(2x)^2 (x)}{(0.15)^2} = 177.8x^3$$

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

$$x = (5.23 \times 10^{-10})^{1/3} = 8.06 \times 10^{-4} \text{ M}$$

$$[\text{H}_2] = 2x = 1.61 \times 10^{-3} \text{ M} \quad [\text{S}_2] = x = 8.06 \times 10^{-4} \text{ M}$$

17.44 $2 \text{ICl}(\text{g}) \rightleftharpoons \text{I}_2(\text{g}) + \text{Cl}_2(\text{g})$, $K_c = 0.110$, $[\text{ICl}]_0 = 0.50 \text{ mol}/5.0 \text{ L} = 0.10 \text{ M}$

	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 H₂S will decompose.

$$K_c = 0.110 = \frac{[\text{I}_2][\text{Cl}_2]}{[\text{ICl}]^2} = \frac{(x)(x)}{(0.10-2x)^2} = \frac{x^2}{(0.10-2x)^2}$$

Take square root of both sides

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

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

$$0.0332 - 0.664x = x$$

$$1.664x = 0.0332$$

$$x = 0.01995 = 0.020$$

$$[\text{I}_2] = [\text{Cl}_2] = x = 0.020 \text{ M} \quad [\text{ICl}] = 0.10 - 2x = 0.10 - 2(0.02) = 0.060 \text{ M}$$

17.46 $4 \text{NH}_3(\text{g}) + 3 \text{O}_2(\text{g}) \rightleftharpoons 2 \text{N}_2(\text{g}) + 6 \text{H}_2\text{O}(\text{g})$
 $[\text{NH}_3]_0 = [\text{O}_2]_0 = 0.015 \text{ mol}/1.00 \text{ L} = 0.015 \text{ M}$, $[\text{N}_2]_{\text{equil}} = 1.96 \times 10^{-3} \text{ M}$

Let's make an ICE Table

	NH ₃	O ₂	N ₂	H ₂ O
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, $[\text{N}_2]$.

$$[\text{N}_2]_{\text{equil}} = 2x = 1.96 \times 10^{-3} \Rightarrow x = 9.80 \times 10^{-4}$$

$$[\text{H}_2\text{O}] = 6x = 5.88 \times 10^{-3}, \quad [\text{NH}_3] = 0.015 - 4x = 1.108 \times 10^{-2}, \quad [\text{O}_2] = 0.015 - 3x = 1.206 \times 10^{-2}$$

$$K_c = \frac{[\text{N}_2]^2 [\text{H}_2\text{O}]^6}{[\text{NH}_3]^4 [\text{O}_2]^3} = \frac{(1.96 \times 10^{-3})^2 (5.88 \times 10^{-3})^6}{(1.108 \times 10^{-2})^4 (1.206 \times 10^{-2})^3} = 6.01 \times 10^{-6} \approx 6.0 \times 10^{-6}$$



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

$$K_p = 0.403 = \frac{P_{\text{CO}_2}}{P_{\text{CO}}} = \frac{x}{1.00 - x}$$

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

$$0.403 - 0.403x = x$$

$$1.403x = 0.403$$

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

$$P_{\text{CO}} = 1.00 - x = 1.00 - 0.287 = 0.713 \text{ atm}, \quad P_{\text{CO}_2} = x = 0.287 \text{ atm}$$

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) $\text{F}_2\text{(g)} \rightleftharpoons 2 \text{F(g)}$: Equil. will move to the right. More F and less F₂. No change in K

(b) $2 \text{CH}_4\text{(g)} \rightleftharpoons \text{C}_2\text{H}_2\text{(g)} + 3 \text{H}_2\text{(g)}$: Equil. will move to the right. More C₂H₂ and H₂ and less CH₄. 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 ($\Delta H^\circ_{\text{rxn}} = -151 \text{ kJ}$). Equilibrium will move to right and K will increase.

(b) Exothermic reaction ($\Delta H^\circ_{\text{rxn}} = -451 \text{ kJ}$). 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.