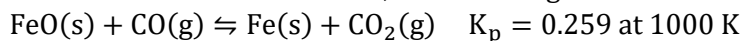


1. When iron(II) and carbon monoxide are mixed, the following reaction occurs



What are the equilibrium partial pressures of CO and CO₂ at 1000 K if the initial partial pressures are P_{CO} = 1.000 atm and P_{CO₂} = 0.500 atm? What is the total pressure at equilibrium?

$$Q_p = \frac{P_{\text{CO}_2}}{P_{\text{CO}}} = \frac{0.500}{1.000} = 0.500$$

Since Q_p is larger than K_p, the reaction will proceed towards the reactants

	FeO(s)	+	CO(g)	⇌	Fe(s)	+	CO ₂ (g)
Initial			1.000 atm				0.500 atm
Change			+x				-x
Equil.			1.000 + x				0.500 - x

$$K_p = \frac{P_{\text{CO}_2}}{P_{\text{CO}}} = 0.259$$

$$0.259 = \frac{0.500 - x}{1.000 + x}$$

$$0.259 + 0.259x = 0.500 - x$$

$$-0.241 = -1.259x$$

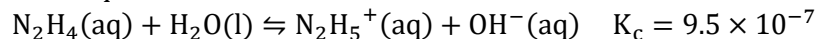
$$x = \frac{-0.241}{-1.259} = 0.191$$

$$P_{\text{CO}} = 1.000 - x = 1.000 - 0.191 = 0.809 \text{ atm}$$

$$P_{\text{CO}_2} = 0.500 + x = 0.500 + 0.191 = 0.691 \text{ atm}$$

$$P_{\text{total}} = 0.809 \text{ atm} + 0.691 \text{ atm} = 1.500 \text{ atm}$$

2. A 0.200 solution of N_2H_4 is mixed with water. What are the concentrations of N_2H_4 , $N_2H_5^+$, and OH^- at equilibrium?



	$N_2H_4(aq)$	+	$H_2O(l)$	\rightleftharpoons	$N_2H_5^+(aq)$	+	$OH^-(aq)$
Initial	0.200 M				0 M		0 M
Change	-x				+x		+x
Equil.	0.200 - x				x		x

$$K_c = \frac{[N_2H_5^+][OH^-]}{[N_2H_4]}$$

$$9.5 \times 10^{-7} = \frac{x^2}{0.200 - x}$$

Assume x is much smaller than 0.200

$$9.5 \times 10^{-7} = \frac{x^2}{0.200}$$

$$1.9 \times 10^{-7} = x^2$$

$$x = 4.4 \times 10^{-4} \text{ M}$$

Check to make sure that the assumption was valid

$$\frac{4.4 \times 10^{-4}}{0.200} \times 100 = 0.22$$

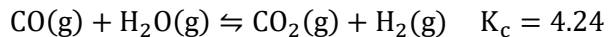
Since $0.22 < 5$, the approximation is valid!

$$[N_2H_4] = 0.200 - x = 0.200 - 4.4 \times 10^{-4} = 0.200 \text{ M}$$

$$[N_2H_5^+] = x = 4.4 \times 10^{-4} \text{ M}$$

$$[OH^-] = x = 4.4 \times 10^{-4} \text{ M}$$

3. Calculate the equilibrium concentrations of CO_2 , H_2 , CO , and H_2O at 800 K if only CO and H_2O are present initially at concentrations of 0.150 M each.



	CO(g)	+	$\text{H}_2\text{O(g)}$	\rightleftharpoons	$\text{CO}_2\text{(g)}$	+	$\text{H}_2\text{(g)}$
<i>Initial</i>	0.150 M		0.150 M		0 M		0 M
<i>Change</i>	-x		-x		+x		+x
<i>Equil.</i>	$0.150 - x$		$0.150 - x$		x		x

$$K_c = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

$$4.24 = \frac{x^2}{(0.150 - x)^2}$$

$$\sqrt{4.24} = \sqrt{\frac{x^2}{(0.150 - x)^2}}$$

$$2.06 = \frac{x}{0.150 - x}$$

$$0.309 - 2.06x = x$$

$$0.309 = 3.06x$$

$$x = 0.101 \text{ M}$$

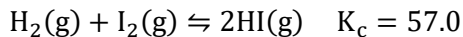
$$[\text{CO}] = 0.150 - x = 0.150 - 0.101 = 0.049 \text{ M}$$

$$[\text{H}_2\text{O}] = 0.150 - x = 0.150 - 0.101 = 0.049 \text{ M}$$

$$[\text{CO}_2] = x = 0.101 \text{ M}$$

$$[\text{H}_2] = x = 0.101 \text{ M}$$

4. Calculate the equilibrium concentrations of H_2 , I_2 , and HI at 700 K if the initial concentrations are $[\text{H}_2] = 0.100 \text{ M}$ and $[\text{I}_2] = 0.200 \text{ M}$.



	$\text{H}_2(\text{g})$	+	$\text{I}_2(\text{g})$	\rightleftharpoons	$2 \text{HI}(\text{g})$
<i>Initial</i>	0.100 M		0.200 M		0 M
<i>Change</i>	-x		-x		+2x
<i>Equil.</i>	$0.100 - x$		$0.200 - x$		2x

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

$$57 = \frac{4x^2}{(0.100 - x)(0.200 - x)}$$

Assume that x is much smaller than 0.100

$$57 = \frac{4x^2}{(0.100)(0.200)}$$

$$1.14 = 4x^2$$

$$0.285 = x^2$$

$$x = 0.534$$

Check to make sure that the assumption was valid

$$\frac{0.534}{0.100} \times 100 = 534$$

Since $534 > 5$, the approximation is invalid and the quadratic equation must be used

$$57 = \frac{4x^2}{(0.100 - x)(0.200 - x)}$$

$$57 = \frac{4x^2}{(0.0200 - 0.300x + x^2)}$$

$$1.14 - 17.1x + 57x^2 = 4x^2$$

$$53x^2 - 17.1x + 1.14 = 0$$

$$\frac{-17.1 \pm \sqrt{17.1^2 - 4(53)(1.14)}}{2(53)}$$

$$\frac{17.1 \pm 7.1}{106} = 0.228 \text{ and } 0.0943$$

Discard the solution that gives 0.228 because the H_2 concentration can't change by more than its initial value (0.100 M)

$$[\text{H}_2] = 0.100 - x = 0.100 - 0.0943 = 0.006 \text{ M}$$

$$[\text{I}_2] = 0.200 - x = 0.200 - 0.0943 = 0.106 \text{ M}$$

$$[\text{HI}] = 2x = 2(0.0943) = 0.189 \text{ M}$$