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

$$FeO(s) + CO(g) \leftrightharpoons Fe(s) + CO_2(g) \quad K_p = 0.259 \text{ at } 1000\text{ K}$$

What are the equilibrium partial pressures of CO and CO_2 at 1000 K if the initial partial pressures are $P_{CO} = 1.000$ atm and $P_{CO_2} = 0.500$ atm? What is the total pressure at equilibrium?

$$Q_p = \frac{P_{CO_2}}{P_{CO}} = \frac{0.500}{1.000} = 0.500$$

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

	FeO(s)	+	CO(g)	$\leftrightharpoons$	Fe(s)	+	$CO_2(g)$
Initial			1.000 atm				0.500 atm
Change			$+x$				$-x$
Equil.			$1.000 + x$				$0.500 - x$

$$K_p = \frac{P_{CO_2}}{P_{CO}} = 0.259$$

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

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

$$-0.241 = -1.259x$$

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

$$P_{CO} = 1.000 - x = 1.000 + 1.191 = 1.191 \text{ atm}$$

$$P_{CO_2} = 0.500 + x = 0.500 - 0.191 = 0.309 \text{ atm}$$

$$P_{total} = 1.191 \text{ atm} + 0.309 \text{ atm} = 1.500 \text{ atm}$$

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?

$$N_2H_4(aq) + H_2O(l) \leftrightharpoons N_2H_5^+(aq) + OH^-(aq) \quad K_c = 9.5 \times 10^{-7}$$

	$N_2H_4(aq)$ +	$H_2O(l)$	$\leftrightharpoons$	$N_2H_5^+(aq)$ +	$OH^-(aq)$
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.

$$CO(g) + H_2O(g) \leftrightharpoons CO_2(g) + H_2(g) \quad K_c = 4.24$$

	$CO(g)$	$+$	$H_2O(g)$	$\leftrightharpoons$	$CO_2(g)$	$+$	$H_2(g)$
Initial	0.150 M		0.150 M		0 M		0 M
Change	$-x$		$-x$		$+x$		$+x$
Equil.	$0.150 - x$		$0.150 - x$		x		x

$$K_c = \frac{[CO_2][H_2]}{[CO][H_2O]}$$

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

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

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

$$0.309 - 2.06x = x$$
$$0.309 = 3.06x$$
$$x = 0.101\ M$$
$$[CO] = 0.150 - x = 0.150 - 0.101 = 0.049\ M$$
$$[H_2O] = 0.150 - x = 0.150 - 0.101 = 0.049\ M$$
$$[CO_2] = x = 0.101\ M$$
$$[H_2] = x = 0.101\ M$$

4. Calculate the equilibrium concentrations of H_2, I_2, and HI at 700 K if the initial concentrations are $[H_2] = 0.100$ M and $[I_2] = 0.200$ M.

$$\mathrm{H_2(g) + I_2(g) \leftrightharpoons 2HI(g)} \quad K_c = 57.0$$

	$H_2(g)$	+	$I_2(g)$	$\leftrightharpoons$	$2\ HI(g)$
Initial	0.100 M		0.200 M		0 M
Change	$-x$		$-x$		$+2x$
Equil.	$0.100 - x$		$0.200 - x$		$2x$

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

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

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)

$$[H_2] = 0.100 - x = 0.100 - 0.0943 = 0.006 \text{ M}$$

$$[I_2] = 0.200 - x = 0.200 - 0.0943 = 0.106 \text{ M}$$

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