1. At a particular temperature, $K_c = 2.0 \times 10^{-6}$ for the reaction

$$2\ CO_2\ (g) \rightleftharpoons 2\ CO\ (g) + O_2\ (g)$$

If 2.0 mol CO_2 is initially placed into a 5.0 L vessel, calculate the equilibrium concentrations of all species.

	$2\ CO_2\ (g)$	$\rightleftharpoons$ $2\ CO\ (g)$	+ $O_2\ (g)$
Initial (*M*)	0.40	0	0
Change (*M*)	$-2x$	$+2x$	$+x$
Equil (*M*)	$0.40 - 2x$	$2x$	x

$$K_c = \frac{[CO]^2[O_2]}{[CO_2]^2}$$

$$2.0 \times 10^{-6} = \frac{[2x]^2[x]}{[0.40 - 2x]^2}$$

Because the value of K_c is very small, very little reactant goes to product. We assume that $0.40 - 2x \approx 0.40$.

$$2.0 \times 10^{-6} = \frac{[2x]^2[x]}{[0.40]^2}$$

$3.2 \times 10^{-7} = 4x^3$
$x^3 = 8.0 \times 10^{-8}$
$x = 4.3 \times 10^{-3}\ M$

$[O_2] = x = 4.3 \times 10^{-3}\ M$
$[CO] = 2x = 8.6 \times 10^{-3}\ M$
$[CO_2] = 0.40 - 2x = 0.39\ M$

Checking the assumption:

$$\frac{2x}{0.40} \times 100 = \frac{8.6 \times 10^{-3}}{0.40} \times 100 = 2.2\% < 5\%$$

The assumption was valid.

2. At 2200°C, $K_p = 0.050$ for the reaction

$$N_2\ (g) + O_2\ (g) \rightleftharpoons 2\ NO\ (g)$$

What is the partial pressure of NO in equilibrium with N_2 and O_2 that was placed in a flask at initial pressures of 0.80 and 0.20 atm, respectively?

	$N_2\ (g)$	+ $O_2\ (g)$	$\rightleftharpoons$ $2\ NO\ (g)$
Initial (pressure)	0.80	0.20	0
Change (pressure)	$-x$	$-x$	$+2x$
Equil (pressure)	$0.80 - x$	$0.20 - x$	$2x$

$$K_p = \frac{P_{NO}^2}{P_{N_2} P_{O_2}}$$

$$0.050 = \frac{(2x)^2}{(0.80 - x)(0.20 - x)}$$

$$0.050 = \frac{4x^2}{0.16 - x + x^2}$$

$$0.0080 - 0.050x + 0.050x^2 = 4x^2$$

$$3.95x^2 + 0.050x - 0.0080 = 0$$

A quadratic equation of the form $ax^2 + bx + c = 0$, has the solutions: $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$x = \frac{-0.050 \pm \sqrt{(0.050)^2 - 4(3.95)(-0.0080)}}{2(3.95)}$$

$$x = \frac{-0.050 \pm \sqrt{0.1289}}{7.90}$$

$$x = \frac{-0.050 \pm 0.359}{7.90}$$

$x = 0.039$ atm or $x = -0.052$ atm

The correct answer for x is 0.039 atm because you cannot have a negative pressure.

$P_{NO} = 2x = 0.078$ atm

3. A 1.00 L flask was filled with 2.00 mol SO_2 (*g*) and 2.00 mol NO_2 (*g*) and heated. After equilibrium was reached, it was found that 1.30 mol NO (*g*) was present. Assume that the reaction

$$SO_3\,(g) + NO\,(g) \rightleftharpoons SO_2\,(g) + NO_2\,(g)$$

occurs under these conditions. Calculate the value of the equilibrium constant, K_c, for the above reaction.

	SO_3 (*g*) +	NO (*g*) ⇌	SO_2 (*g*) +	NO_2 (*g*)
Initial (*M*)	2.00	2.00	0	0
Change (*M*)	$-x$	$-x$	$+x$	$+x$
Equil (*M*)	$2.00 - x$	$2.00 - x$	x	x

$$K_c = \frac{[SO_2][NO_2]}{[SO_3][NO]}$$

$$K_c = \frac{x^2}{(2.00 - x)^2}$$

Since the problem asks you to solve for K_c, it must indicate in the problem what the value of x is. The concentration of NO at equilibrium is given to be 1.30 *M*. In the table above, we have the concentration of NO set equal to $2.00 - x$.

$2.00 - x = 1.30$
$x = 0.70$

Substituting back into the equilibrium constant expression:

$$K_c = \frac{x^2}{(2.00 - x)^2}$$

$$K_c = \frac{(0.70)^2}{(2.00 - 0.70)^2}$$

$K_c = 0.290$