1. For the following system, 2.000 moles of CO_2 is placed in a 5.000-L flask and allowed to reach equilibrium, at a particular temperature. The value of K_c for this reaction is 2.000×10^{-6}. What are the concentrations of all substances at equilibrium?

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

Answer: $[CO_2] = 0.3914$ M $[CO] = 8.618 \times 10^{-3}$ M $[O_2] = 4.309 \times 10^{-3}$ M

2. For the following reaction the equilibrium constant, K_p, is 1.00×10^2. Suppose a mixture consists of 0.5000 atm of C, 0.01000 atm of A and 0.01000 atm of B. What are the partial pressures of all substances at equilibrium? (HINT: You must first determine which direction the reaction proceeds.)

$$A(g) + B(g) \rightleftarrows 2\ C(g)$$

Answer: $P_A = 0.0433$ atm $P_B = 0.0433$ atm $P_C = 0.4333$ atm

3. For the following reaction the equilibrium constant, K_c, is 3.000×10^{-5}. Suppose a mixture consists of 0.2000 moles of A and 0.02000 moles of B (and NO C) in a 2.000-L flask. What are the concentrations of all substances at equilibrium?

$$A(aq) \rightleftarrows B(aq) + C(aq)$$

Answer: $[A] = 0.0997$ M $[B] = 1.030 \times 10^{-2}$ M $[C] = 3.000 \times 10^{-4}$ M

4. For the following system, 2.000 moles of A is placed in a 1.000-L flask and allowed to reach equilibrium, at a particular temperature. The value of K_c for this reaction is 4.000×10^3. What are the concentrations of all substances at equilibrium?

$$2\ A(g) \rightleftarrows 2\ B(g) + C(g)$$

Answer: $[A] = 3.162 \times 10^{-2}$ M $[B] = 1.968$ M $[C] = 0.984$ M

1. For the following system, 2.000 moles of CO_2 is placed in a 5.000-L flask and allowed to reach equilibrium, at a particular temperature. The value of K_c for this reaction is 2.000×10^{-6}. What are the concentrations of all substances at equilibrium?

$[CO_2]_0 = 2.000$ mol/ 5.000 L = 0.4000 M

	$2\,CO_2(g)$	$\rightleftarrows$	$2\,CO(g)$	+	$O_2(g)$	(Done using Molarity)
I	0.4000		0		0	
C	- 2x		+ 2x		+ x	
E	0.4000 -2x		2x		x	

$$K_c = \frac{[CO]^2\,[O_2]}{[CO_2]^2} = \frac{(2x)^2\,(x)}{(0.4000 - 2x)^2} = \frac{4\,(x^3)}{(0.4000 - 2x)^2} = 2.000 \times 10^{-6}$$

This is a cubic equation which you can't solve easily with your calculators. However, K_c is pretty small so there's mostly reactants (very little product) at equilibrium. Not much of the CO_2 reacts. We can assume (we will check the assumption later) $2x << 0.4000$ and $0.4000 - 2x \approx 0.4000$.

$$\frac{4\,(x^3)}{(0.4000)^2} = 2.000 \times 10^{-6}$$ This can be easily solved for "x".

$x = 4.30\underline{\mathbf{8}}8 \times 10^{-3}$

Answer:

$[CO_2] = 0.4000 - 2(0.00430\underline{\mathbf{8}}8) = 0.3914$ M
$[CO] = 2(4.30\underline{\mathbf{8}}8 \times 10^{-3}) = 8.618 \times 10^{-3}$ M
$[O_2] = 4.30\underline{\mathbf{8}}8 \times 10^{-3} = 4.309 \times 10^{-3}$ M

Check % error (in ignoring "2x" compared to 0.4000)

$$\%\ \text{error} = \frac{(0.4000 - 0.3914)}{0.4000} * 100\% = 2.154\ \%\ \text{error}$$ Less than 5% error so the assumption is okay.

Plug these back into the expression for K_c to see if you get the value for K_c given in the problem as a check to see if you've done the problem correctly.

$$K_c = \frac{[CO]^2\,[O_2]}{[CO_2]^2} = \frac{(8.618 \times 10^{-3})^2\,(4.309 \times 10^{-3})}{(0.3914)^2} = 2.0890 \times 10^{-6}$$ (close enough)

A more exact answer obtained by using the method of successive approximations gives:

$[CO_2] = 0.4000 - 2(0.00424\underline{7}6) = 0.3915$ M
$[CO] = 2(4.24\underline{7}6 \times 10^{-3}) = 8.495 \times 10^{-3}$ M
$[O_2] = 4.24\underline{\mathbf{7}}6 \times 10^{-3} = 4.248 \times 10^{-3}$ M

I would be looking for the first set of answers given above (not these more exact answers).

2. For the following reaction the equilibrium constant, K_p, is 1.00 x 10^2. Suppose a mixture consists of 0.5000 atm of C, 0.01000 atm of A and 0.01000 atm of B. What are the partial pressures of all substances at equilibrium? (HINT: You must first determine which direction the reaction proceeds.)

$$A(g) + B(g) \rightleftarrows 2\,C(g)$$

$P_{A,o}$ = 0.01000 atm $P_{B,o}$ = 0.01000 atm $P_{C,o}$ = 0.5000 atm starting partial pressures

We are starting with all the substances so we first have to determine if we are at equilibrium and if not which direction the reaction is proceeding to reach equilibrium (so we know where to put the "-" in the ICE table).

$$Q_p = \frac{(P_{C,o})^2}{(P_{A,o})\,(P_{B,o})} = \frac{(0.5000)^2}{(0.01000)(0.01000)} = 2500$$

$$\begin{array}{ccc} Q & > & K \\ 2500 & > & 100 \end{array}$$

Since Q (2500) > K (100) it means there's too much product and not enough reactant to be at equilibrium. Thus the reaction must proceed from right to left (in reverse direction) to reach equilibrium (the "-" signs in the ICE table must be on the product side).

	A(g)	+	B(g)	⇄	2 C(g)	(Done using atm)
I	0.01000		0.01000		0.5000	
C	+ x		+ x		- 2x	
E	0.01000 + x		0.01000 + x		0.5000 - 2x	

$$K_p = \frac{(P_C)^2}{(P_A)\,(P_B)} = \frac{(0.5000 - 2x)^2}{(0.01000 + x)^2} = 10\underline{\mathbf{0}}$$

The left side is a perfect square so we can solve this by simply taking the square root of both sides.

$$\frac{(0.5000 - 2x)}{(0.01000 + x)} = 10.0$$

$0.5000 - 2x = 10.0\,(0.01000 + x)$

$0.5000 - 2x = 0.1000 + 10.0(x)$

$0.4000 = 12.0\,(x)$

$x = 0.033\underline{\mathbf{3}}33$

Answer: P_A = 0.0433 atm P_B = 0.0433 atm P_C = 0.4333 atm

$P_A = 0.01000 + 0.033\underline{\mathbf{3}}33 = 0.0433$ atm

$P_B = 0.01000 + 0.033\underline{\mathbf{3}}33 = 0.0433$ atm

$P_C = 0.5000 - 2(0.033\underline{\mathbf{3}}33) = 0.4333$ atm

3. For the following reaction the equilibrium constant, K_c, is 3.000×10^{-5}. Suppose a mixture consists of 0.2000 moles of A and 0.02000 moles of B (and NO C) in a 2.000-L flask. What are the concentrations of all substances at equilibrium?

$$A(aq) \rightleftarrows B(aq) + C(aq)$$

$[A]_0$ = 0.2000 mol/2.000 L = 0.1000 M $[B]_0$ = 0.0200 mol/2.000 L = 0.01000 M

	A(aq)	⇄	B(aq)	+	C(aq)
I	0.1000		0.01000		0
C	- x		+ x		+ x
E	0.1000 - x		0.01000 + x		+ x

$$K_c = \frac{[B]\,[C]}{[A]} = \frac{(0.01000 + x)\,(x)}{(0.1000 - x)} = 3.000 \times 10^{-5}$$

This is a quadratic equation which you can solve using the quadratic solution. However, K_c is pretty small so there's mostly reactants (very little product) at equilibrium. Not much of the A reacts. We can assume (we will check the assumption later) "x" is small so x << 0.1000 and x << 0.01000 so 0.1000 - x ≈ 0.1000 and 0.01000 + x ≈ 0.01000.

$$\frac{(0.01000)\,(x)}{0.1000} = 3.000 \times 10^{-5}$$

This can be easily solved for "x".

$$x = 3.000 \times 10^{-4}$$

Need to check % error. We can do that for the assumption in the numerator, that "x" is small compared to 0.01000. If this is okay then it will be okay to ignore the "x" compared to 0.1000.

0.01000 + 0.0003000 = 0.01030 M

$$\%\ error = \frac{(0.01030 - 0.01000)}{0.01000} * 100\% = 3.00\ \%\ error$$

Less than 5% error so the assumption is okay.

Answer: [A] = 0.0997 M [B] = 1.030×10^{-2} M [C] = 3.000×10^{-4} M

[A] = 0.1000 - 0.0003000 = 0.0997 M
[B] = 0.01000 + 0.0003000 = 0.01030 M
[C] = 3.000×10^{-4} M

Solving a quadratic rather than making the assumptions above you get essentially the same answers. For "x" you would get 2.991×10^{-4} (3.000×10^{-4} is about a 0.3 % error).

4. For the following system, 2.000 moles of A is placed in a 1.000-L flask and allowed to reach equilibrium, at a particular temperature. The value of K_c for this reaction is 4.000×10^3. What are the concentrations of all substances at equilibrium?

$[A]_0 = 2.000 \text{ mol}/ 1.000 \text{ L} = 2.000 \text{ M}$

(Done using Molarity)

	2 A(g)	⇄	2 B(g)	+	C(g)
I	2.000		0		0
C	- 2x		+ 2x		+ x
E	2.000 -2x		2x		x

$$K_c = \frac{[B]^2 [C]}{[A]^2} = \frac{(2x)^2 (x)}{(2.000 - 2x)^2} = \frac{4 (x^3)}{(2.000 - 2x)^2} = 4.000 \times 10^3$$

This is a cubic equation which you can't solve easily with your calculators. However, K_c is pretty large so there's mostly products (very little reactant) at equilibrium. The "trick" is to first assume all the reactant reacts to produce nothing but products. Then reverse the reaction and determine how much of the reactants (previously the products in the original reaction) come apart.

Do the first part in which ALL the reactant is converted to the products (reaction goes to completion). Use stoichiometry to determine the concentration of the products, assuming the reaction goes to completion.

(Done using Molarity)

	2 A(g)	⇄	2 B(g)	+	C(g)
I	2.000		0		0
C	- 2.000		+ 2.000		+ 1.000
E	0		2.000		1.000

Now reverse the reaction and determine how much of the B and C react to give A. When you do this the K for this reaction is the inverse of the original K (remember, $K_{rev} = 1/K_{for}$).

	2 B(g)	+	C(g)	⇄	2 A(g)
I	2.000		1.000		0
C	- 2 x		- x		+ 2x
E	2.000 - 2x		1.000 - x		2x

$$\frac{(2x)^2}{(2.000 - 2x)^2 (1.000 - x)} = 2.500 \times 10^{-4}$$

This K is small so this reaction does not proceed very far to the right. This means very little reaction occurs so very little of the B and C react. We can assume (we will check the assumption later) "x" is small so $x << 1.000$ and $2x << 2.000$ so $2.000 - 2x \approx 2.000$ and $1.000 - x \approx 1.000$

***** continued on next page *****

4. (Cont.)

$$\frac{(2x)^2}{(2.000)^2\,(1.000)} = 2.500 \times 10^{-4}$$

This can be easily solved for "x".

$x = 1.58\underline{\mathbf{1}}1 \times 10^{-2}$ M

Answer:

$[A] = 2(0.0158\underline{\mathbf{1}}1) = 0.0316\underline{\mathbf{2}}27$ M $= 3.162 \times 10^{-2}$ M
$[B] = 2.000 - 2(0.0158\underline{\mathbf{1}}1) = 1.96\underline{\mathbf{8}}37$ M $= 1.968$ M
$[C] = 1.000 - 0.0158\underline{\mathbf{1}}1 = 0.98\underline{\mathbf{4}}188$ M $= 0.984$ M

Check % error (in ignoring "x" compared to 1.000)

$$\% \text{ error} = \frac{(1.000 - 0.98\underline{\mathbf{4}}188)}{1.000} * 100\% = 1.58\ \% \text{ error}$$

Less than 5% error so the assumption is okay.

Plug these back into the original expression for K_c to see if you get the value for K_c given in the problem as a check to see if you've done the problem correctly.

$$K_c = \frac{[B]^2\,[C]}{[A]^2} = \frac{(1.96\underline{\mathbf{8}}37)^2\,(0.98\underline{\mathbf{4}}188)}{(0.0316\underline{\mathbf{2}}27)^2} = 3.8\underline{\mathbf{1}}32 \times 10^3$$

This is a 4.7% error from the original K_c.

A more exact answer obtained by using the method of successive approximations gives:

$x = 1.5\underline{\mathbf{4}}46 \times 10^{-2}$

$[A] = 2(0.015\underline{\mathbf{4}}46) = 3.0\underline{\mathbf{8}}929 \times 10^{-2}$ M $= 3.09 \times 10^{-2}$ M
$[B] = 2.000 - 2(0.015\underline{\mathbf{4}}46) = 1.96\underline{\mathbf{9}}10$ M $= 1.969$ M
$[C] = 1.000 - 0.015\underline{\mathbf{4}}46 = 0.98\underline{\mathbf{4}}55$ M $= 0.985$ M

When plugged back into the original expression for K_c one gets 4.0000025×10^3

I would be expecting the first set of answers given above (not these more exact answers) since the % error was less than 5%.