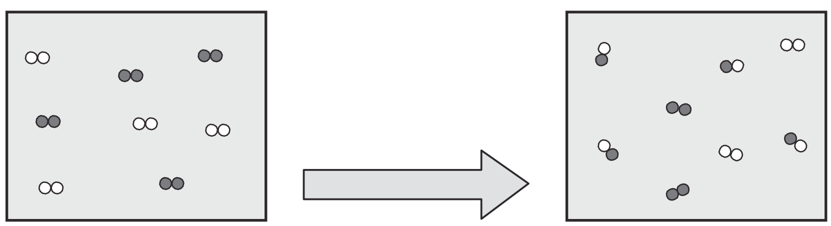
Calculating the Equilibrium Concentration of All Species

Name ______________________ Sec ____________

1. Below are 1.0 L containers for the initial and equilibrium concdition for the reaction

$$A_2(g) + D_2(g) \rightleftharpoons 2AD(g)$$

Calculate the magnitude of the equilibrium constant for the reaction.



Container Before Reaction Started

Container After Reaction Achieves Equilibrium

Show work:

$$K_c = \frac{[AD]^2}{[A_2][D_2]}$$

$$K_c = \frac{(4)^2}{(2)(2)} = 4$$

2. Set up the ICE table for the following general chemical equation. (Assume the reaction proceeds from left to right to establish equilibrium.

$$A_2(g) + D_2(g) \rightleftharpoons 2AD(g)$$

	$A_2(g)$ +	$D_2(g)$ $\rightleftharpoons$	$2AD(g)$	
Initial	$[A_2]_0$	$[D_2]_0$	$[AD]_0$	
Change	$-x$	$-x$	$+2x$	$x = [C_2]_{reacting}$
Equilibrium	$[A_2]_0 - x$	$[D_2]_0 - x$	$[AD]_0 + 2x$	

$$K_c = \frac{[AD]^2}{[A_2][D_2]}$$

$$K_c = \frac{([AD]_0 + 2x)^2}{([A_2]_0 - x)^1([D_2]_0 - x)^1}$$

3. The equilibrium constant, K_p, for the reaction

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$

is 0.0202 at a particular temperature. If the initial partial pressure of H_2 = $[I_2]$ = 0.350 atm, calculate the equilibrium partial pressures of all species

	2HI $\rightleftharpoons$	**H_2**	**+**	**I_2**	
initial	**0**	**0.350 atm**		**0.350 atm**	
change	**+2x**	**−x**		**−x**	**$x = [H_2]_{reacting}$**
equilibrium	**0 + 2x**	**0.350 - x**		**0.350 – x**	

$$K_p = 0.0202 = \frac{P(H_2)\cdot P(I_2)}{P^2(HI)} = \frac{(.350 - x)^2}{(2x)^2}$$

Take the square root of both sides

$$0.142 = \frac{(.350 - x)}{(2x)}$$

$2x \cdot 0.142 = 0.350 - x$
$1.284x = 0.350$
$x = 0.273$ atm
$P(HI) = 2\cdot 0.273$ atm = 0.546 atm
$P(H_2) = P(I_2) = 0.350$ atm – 0.273 atm = 0.077 atm

4. The equilibrium constant, K_c, for the reaction

$$PCl_5 \rightleftharpoons PCl_3 + Cl_2$$

Is 33.3 at 760 ˚C. If 0.400 mol of PCl_5 are placed in a 2.00 Liter container, calculate the equilibrium concentration of all species.

$$[PCl_5]_0 = \frac{0.400\ \text{mol}}{2.00\ \text{L}} = 0.200\ \text{M}$$

	PCl_5 $\rightleftharpoons$	PCl_3	+	Cl_2	
initial	0.200 M	0		0	
change	–x	+x		+x	x = $[PCl_5]_{reacting}$
equilibrium	0.200 – x	0 + x		0 + x	

$$K_p = 33.3 = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{x^2}{0.200\text{ M} - x}$$

Must simplify to a quadratic expression,

$$33.3 = \frac{x^2}{0.200\text{ M} - x}$$

$$33.3\,(0.200\text{ M} - x) = x^2$$

$$6.66 - 33.3x = x^2$$

$$x^2 + 33.3x - 6.66 = 0$$

solving the quadratic equation $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$x = \frac{-33.3 \pm \sqrt{(33.3)^2 - 4(1)(-6.66)}}{2(1)}$$

$$x = \frac{-33.3 \pm 33.698}{2}$$

Use only the positive root

$x = 0.199$ M

$[PCl_3] = [Cl_2] = x = 0.199$ M

$[PCl_5] = 0.200\text{ M} - x = 0.200\text{ M} - 0.199\text{ M} = 0.001$ M

5. The equilibrium constant, K_c, for the reaction

$$PCl_5 \rightleftharpoons PCl_3 + Cl_2$$

Is 33.3 at 760 ˚C. If 0.400 mol of PCl_5 and 1.00 mol of Cl_2 are placed in a 2.00 Liter container, calculate the equilibrium concentration of all species.

$$[PCl_5]_0 = \frac{0.400\text{ mol}}{2.00\text{ L}} = 0.200\text{ M} \qquad [Cl_2]_0 = \frac{1.00\text{ mol}}{2.00\text{ L}} = 0.500\text{ M}$$

	PCl_5 $\rightleftharpoons$	PCl_3	+	Cl_2	
initial	0.200 M	0		0.500	
change	–x	+x		+x	x = $[PCl_5]_{reacting}$
equilibrium	0.200 – x	0 + x		0.500 + x	

$$K_p = 33.3 = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{x\cdot(0.500 + x)}{0.200\ M - x}$$

Must simplify to a quadratic expression,

$$33.3 = \frac{x\cdot(0.500 + x)}{0.200\ M - x}$$

$$33.3\ (0.200\ M - x) = 0.500x + x^2$$

$$6.66 - 33.3x = 0.500x + x^2$$

$$x^2 + 33.8x - 6.66 = 0$$

solving the quadratic equation $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$x = \frac{-33.8 \pm \sqrt{(33.8)^2 - 4(1)(-6.66)}}{2(1)}$$

$$x = \frac{-33.8 \pm 34.192}{2}$$

Use only the positive root

$x = 0.196\ M$

$[PCl_3] = 0.196\ M$

$[Cl_2] = 0.500 + x = 0.500 + 0.196\ M = 0.696$

$[PCl_5] = 0.200\ M - x = 0.200\ M - 0.196\ M = 0.004\ M$