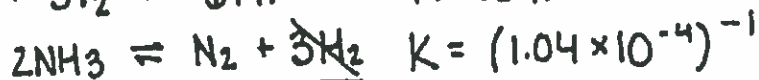
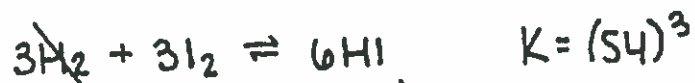


1. Find K_c for the reaction: $2\text{NH}_3(\text{g}) + 3\text{I}_2(\text{g}) \rightleftharpoons 6\text{HI}(\text{g}) + \text{N}_2(\text{g})$, given the information below.



$$\therefore K = 1.5 \times 10^9$$

2. Write an expression for the equilibrium constant of each of the following reactions:



$$K_c = \frac{[\text{C}_2\text{H}_6]^2 [\text{O}_2]}{[\text{C}_2\text{H}_4]^2 [\text{H}_2\text{O}]^2} \quad K_p = \frac{(P_{\text{C}_2\text{H}_6})^2 P_{\text{O}_2}}{(P_{\text{C}_2\text{H}_4})^2 (P_{\text{H}_2\text{O}})^2}$$



~~ignore solids & liquids~~

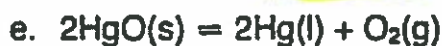
$$K_c = [\text{Mg}^{2+}] [\text{OH}^{-}]^2$$



$$K_c = \frac{[\text{NH}_4^{+}] [\text{OH}^{-}]}{[\text{NH}_3]}$$



$$K_c = \frac{[\text{H}^{+}] [\text{HCOO}^{-}]}{[\text{HCOOH}]}$$



$$K_c = [\text{O}_2]$$

3. The following equilibrium constants were determined at 1123 K:



Calculate the equilibrium constant for the following reaction at 1123 K:



$$K_1 = \frac{[\text{CO}]^2}{[\text{CO}_2]} = 1.3 \times 10^{14}$$

$$K_2 = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]} = 6.0 \times 10^{-3}$$

$$K_3 = \frac{[\text{COCl}_2]^2}{[\text{CO}_2][\text{Cl}_2]^2} = K_1(K_2)^2$$

$$\therefore K_3 = 4.7 \times 10^9$$

4. The equilibrium constant, K_c , for the reaction $\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightleftharpoons 2\text{HF(g)}$ has the value 2.1×10^3 at a particular temperature. When the system is analyzed at equilibrium at this temperature, the concentrations of $\text{H}_2(\text{g})$ and $\text{F}_2(\text{g})$ are both found to be 0.0021 M. What is the concentration of HF(g) in the equilibrium system under these conditions?

$$K_c = \frac{[\text{HF}]^2}{[\text{H}_2][\text{F}_2]} \Rightarrow 2.1 \times 10^3 = \frac{[\text{HF}]^2}{(0.0021)(0.0021)}$$

$$[\text{HF}] = 0.096 \text{ M}$$

5.

At 25°C, $K_p = 5.3 \times 10^5$ for the reaction

When a certain partial pressure of $\text{NH}_3(\text{g})$ is put into an otherwise empty rigid vessel at 25°C, equilibrium is reached when 50.0% of the original ammonia has decomposed. What was the original partial pressure of ammonia before any decomposition occurred?

$$K_p = \frac{(P_{\text{NH}_3})^2}{(P_{\text{N}_2})(P_{\text{H}_2})^3}$$

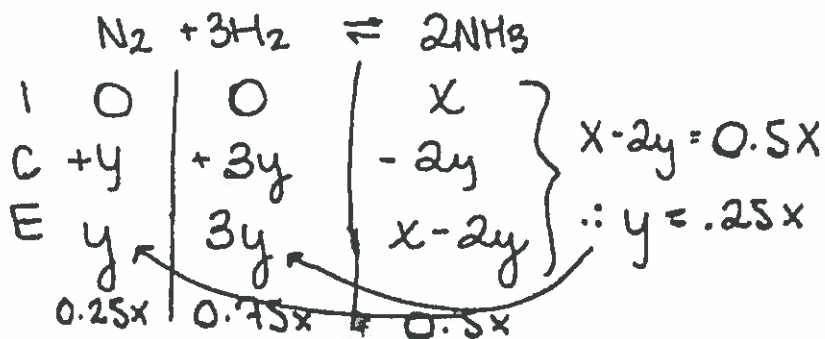
initial pressures:

$$P_{\text{NH}_3} = X$$

$$P_{\text{H}_2} = 0$$

$$P_{\text{N}_2} = 0$$

\therefore reaction proceeds left



$$X - 2y = 0.5X$$

$$\therefore y = .25x$$

$$5.3 \times 10^5 = \frac{(0.5x)^2}{(0.25x)(0.75x)^3}$$

$$5.3 \times 10^5 = 2.37037/x^2$$

$$\therefore X = 2.1 \times 10^{-3} \text{ atm} = P_{\text{NH}_3}(\text{initial})$$

6.

For the following reaction:



the equilibrium constant, K_c , has been determined to be 6.0×10^{-3} at 1123 K. If a mixture of 3.0 M CO, 2.0 M Cl_2 and 0.25 M COCl_2 is put in a vessel, which way will the reaction proceed to reach equilibrium?

$$Q = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]} = \frac{(0.25\text{M})}{(3.0\text{M})(2.0\text{M})} = 0.0417$$

$Q > K \therefore$ reaction proceeds left

7. A mixture of 0.500 mol H_2 and 0.500 mol I_2 was placed in a 1.00 L steel container at $430^\circ C$. The equilibrium constant, K_c for the reaction
- $$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$
- is 5.43×10^{-5} . Calculate the equilibrium concentrations of all components.

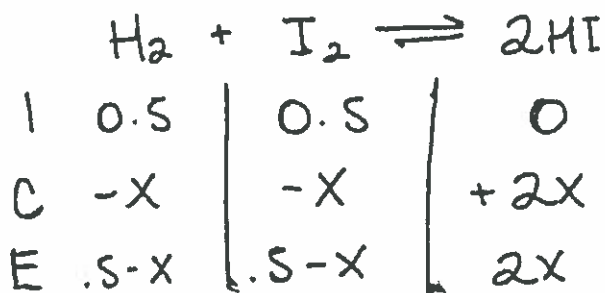
$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

initial conc:

$$[HI]_0 = 0$$

$$[H_2]_0 = 0.5 M$$

$$[I_2]_0 = 0.5 M$$



$$5.43 \times 10^{-5} = \frac{(2x)^2}{(.5-x)(.5-x)}$$

\therefore reaction proceeds right

* take square root of both sides

$$7.369 \times 10^{-3} = 2x / (.5-x)$$

$$-7.369 \times 10^{-3}x + 3.684 \times 10^{-3} = 2x$$

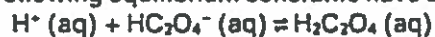
$$X = 1.835 \times 10^{-3}$$

$$[HI] = 3.67 \times 10^{-3} M$$

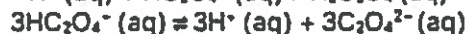
$$[H_2] = 0.498 M$$

$$[I_2] = 0.498 M$$

8. The following equilibrium constants have been determined for oxalic acid at $25^\circ C$:

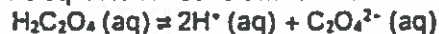


$$K_1 = 15.384$$



$$K_2 = 2.27 \times 10^{-13}$$

Calculate the equilibrium constant for the following reaction at the same temperature:



$$K_3 = ???$$

flip equation 1:



divide equation 2:



$$K_c = (15.384)^{-1} (2.27 \times 10^{-13})^{1/3} = 3.965 \times 10^{-6}$$

9.

For the following reaction:



The initial concentration of PCl_5 is 0.200 moles per liter and there are no products in the system when the reaction starts. If the equilibrium constant is 0.030, calculate all the concentrations at equilibrium.

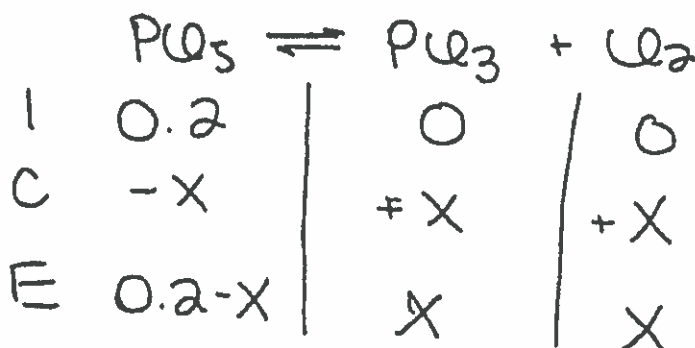
$$K_c = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$

$$[\text{PCl}_5]_0 = 0.2 \text{ M}$$

$$[\text{PCl}_3]_0 = 0$$

$$[\text{Cl}_2]_0 = 0$$

\therefore proceeds right



$$0.030 = \frac{x^2}{0.2-x}$$

$$0.006 - 0.03x = x^2$$

$$0 = x^2 + 0.03x - 0.006$$

using quadratic formula...

$$x = 0.063899 \text{ \& } - 0.0939$$

$$[\text{PCl}_5] = 0.136 \text{ M}$$

$$[\text{PCl}_3] = 0.0639 \text{ M}$$

$$[\text{Cl}_2] = 0.0639 \text{ M}$$

10.

Given this equation:



2.00 g of COCl_2 and 5.00 g of Cl_2 are placed in a 2.50 L flask. Calculate all three equilibrium partial pressures when $K_p = 0.680$ at 25°C

$$K_p = \frac{P_{\text{CO}} \cdot P_{\text{Cl}_2}}{P_{\text{COCl}_2}}$$

Initial Pressures:

$$P_{\text{CO}} = 0 \because \text{rxn proceeds right}$$

$$P_{\text{Cl}_2} = \frac{nRT}{V} = \frac{5.00\text{g} \left(\frac{\text{mol}}{70.5\text{g}}\right) (0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}) (298.15\text{K})}{2.50\text{L}}$$

$$= 0.694\text{M}$$

$$P_{\text{COCl}_2} = \frac{2.00\text{g} \left(\frac{\text{mol}}{98.51\text{g}}\right) (0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}) (298.15\text{K})}{2.50\text{L}}$$

$$= 0.1987\text{M}$$



I	.1987		0		.694
C	-X		+X		+X
E	.1987-X		X		.694+X

$$[\text{Cl}_2] = 0.786\text{M}$$

$$[\text{CO}] = 0.0919\text{M}$$

$$[\text{COCl}_2] = 0.107\text{M}$$

$$0.68 = \frac{X(.694+X)}{.1987-X}$$

$$.68(.1987-X) = X(0.694+X)$$

$$.1351 - .68X = .694X + X^2$$

$$0 = X^2 + 1.378X - 0.1351$$

$$X = 0.09191$$