General Chemistry II
PLI Sheet 6 AK
July 19, 2021

1. **Write the expression for K_C for the following reactions. In each case indicate whether the reaction is homogeneous or heterogeneous.**
   a. **$Ni(CO)_4$ (g) ↔ Ni (s) + 4 CO (g) (1)**

      $K_C = [CO]^4 / [Ni(CO)_4]$, heterogeneous

   b. **H_2O (l) ↔ H^+ (aq) + OH^- (aq) (1)**

      $K_C = [H^+][OH^-]$, homogeneous

   c. **2Ag (s) + Zn^{2+} (aq) ↔ 2 Ag^+ (aq) + Zn(s) (1)**

      $K_C = [Ag^+]^2 / [Zn^{2+}]$, heterogeneous

2. **Consider the following reaction: Cl_2 (g) + I_2 (g) ↔ 2ICl (g) $K_p = 81.9$ at 25℃. A reaction mixture initially at 25℃ initially contains $P_{I2} = 0.100$ atm, $P_{Cl2} = 0.100$ atm, $P_{ICl} = 0.100$ atm. Find the equilibrium partial pressures of I_2, Cl_2, and ICl at this temperature.**

First, you must determine which direction the reaction will proceed in. To do this, calculate Q_p and compare it to K_p.

$$Q_p = \frac{P_{ICl}^2}{P_{Cl_2} * P_{I_2}} = \frac{(0.100)^2}{(0.100)(0.100)} = 1$$

Since Q_p is less than K_p, this means not enough product has formed to be at equilibrium, so the reaction will shift right to create more product accordingly. Set up an ICE chart!

	P_{Cl2}	P_{I2}	P_{ICl}
I	0.100	0.100	0.100
C	-x	-x	+2x
E	0.100-x	0.100-x	0.100+2x

$$K_p = \frac{P_{ICl_{eq}}^2}{P_{Cl_{2eq}} * P_{I_{2eq}}} = \frac{(0.100 + 2x)^2}{(0.100 - x)(0.100 - x)} = 81.9$$

$$\frac{(0.100 + 2x)^2}{(0.100 - x)^2} = 81.9$$

$$\frac{\sqrt{(0.100 + 2x)^2}}{\sqrt{(0.100 - x)^2}} = \sqrt{81.9}$$

$$\frac{0.100 + 2x}{0.100 - x} = 9.05$$

$$0.100 + 2x = 0.905 - 9.05x$$

$$11.05x = 0.805 \rightarrow x = 0.0729\ atm$$

Now, you can calculate the equilibrium partial pressures

$$P_{Cl_{2eq}} = 0.100 - x = 0.100\ atm - 0.0729\ atm = 0.0271\ atm$$

$$P_{I_{2eq}} = 0.100 - x = 0.100\ atm - 0.0729\ atm = 0.0271\ atm$$

$$P_{ICl_{eq}} = 0.100\ atm + 2(0.0729\ atm) = 0.246\ atm$$

3. **Consider the following reaction: NiO (s) + CO (g) ↔ Ni (s) + CO_2 (g) $K_c = 4.0 \times 10^3$ at 1500 K. If a mixture of solid nickel(II) oxide and 0.20 M carbon monoxide is allowed to come to equilibrium at 1500 K, what will be the equilibrium concentration of CO_2?**

First, it is important to think about what should be part of the equilibrium expression. Remember that solids are not included, so only CO and CO_2 will be part of the equilibrium expression. Therefore,

$$K_c = \frac{[CO_2]_{eq}}{[CO]_{eq}}$$

Set up an ICE chart.

	[CO]	[CO_2]
I	0.20	0
C	-x	+x
E	0.20 - x	x

$$K_c = 4.0 * 10^3 = \frac{[CO_2]_{eq}}{[CO]_{eq}} = \frac{x}{0.20 - x}$$

$$x = 800 - 4.0 * 10^3 x$$

$$4001x = 800$$

$$x = 0.20 = [CO_2]_{eq}$$

The large equilibrium constant tells you that the reaction is very product favored, which explains why the [CO_2] at equilibrium is nearly 0.20 M (the initial [CO]).

4. **Consider the reaction: CO (g) + 2H_2 (g) ↔ CH_3OH (g)**
A reaction mixture at 780℃ initially contains [CO] = 0.500 M and [H_2] = 1.00 M. At equilibrium, the CO concentration is found **Text to be 0.150 M. What is the value of the equilibrium constant?**

	[CO] +	[H_2]	⇔ [CH_3OH]
I	0.500	1.00	0

C	-x	-2x	+x
E	0.500 – x	1.00-2x	x

The equilibrium CO concentration is found to be 0.150 M, so then we can solve for x.

$$0.500\ M - x = 0.150\ M$$

$$x = 0.350\ M$$

Therefore,

$$[H_2]_{eq} = 1.00\ M - 2(0.350\ M) = 0.300\ M$$

$$[CH_3OH]_{eq} = 0.350\ M$$

Now, to calculate the equilibrium constant.

$$K_c = \frac{[CH_3OH]_{eq}}{[CO]_{eq}[H_2]_{eq}^2} = \frac{0.350\ M}{(0.150\ M)(0.300\ M)^2} = 25.93$$

5. **Consider the chemical equation and equilibrium constant for the synthesis of NH_3 at 25℃: N_2 (g) + $3H_2$ (g) ↔ $2NH_3$ (g) K_{c1} = 3.7 x 10^8**

Calculate the equilibrium constant for the following reaction at 25℃:
NH_3 (g) ↔ ½ N_2 (g) + 3/2 H_2 (g)

You should first notice that the reaction you need to calculate the equilibrium constant for is the reverse of the first reaction. Because K_C takes the general form of [$\text{product}^{\text{coefficient}}$] / [$\text{reactants}^{\text{coefficient}}$], it means that for this new equation, the first step to finding the new K_C is to take the reciprocal of the known $K_{C.}$ This intuitively makes sense because the old reactants are the new products and the old products are the new reactants; thus, the K_C equation has more or less just been flipped.

So, we now have a preliminary value of 2.7 x 10^{-9}

But you should also notice that the coefficients of the reverse reaction have been changed. It looks like each coefficient has been divided by 2. How should we manipulate the preliminary value we have?

We know by the law of mass action that the coefficients of the species involved in the reaction become exponents when calculating K_C. So...that would mean that each of the species has now had “½” added to their exponent.

For example, $[H_2]^3$ in the original equation becomes $[H_2]^{3/2}$. Since the only operations involved in calculating K_C are multiplication and division, his is the same as taking the square root of our preliminary value.

sqrt(2.7×10^{-9}) = **5.2×10^{-5}**

You can do these manipulations in any order!!! You can do the square root before taking the reciprocal! Try it yourself :)