Chemistry 330

Name: **Date:**

§ 15.03a Equilibrium: Calculating Equilibrium Concentrations

Most Important Idea(s):

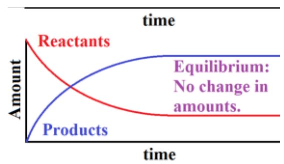
Purpose

The purpose of this activity is to determine the equilibrium concentrations in a chemical reaction.

Background

A dynamic equilibrium in a chemical reaction exists when the rate of the forward reaction (k_1) equals the rate of the reverse reaction (k_{-1}):

$$aA + bB \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} cC + dD$$



The general equation for equilibrium for the reaction above is:

$$\boxed{K = \frac{[C]^c[D]^d}{[A]^a\,[B]^b}}$$

Equilibrium calculations will give you:

1. *The equilibrium constant (K = general constant; Kc = specific for concentrations, expressed in molarity; Kp = specific for gas pressures, which equals the amount of moles according to Avogadro's law), and*
2. *The concentrations of reactants and products at equilibrium.*

Model 1: Calculating the Equilibrium Constant, Kc.

The equilibrium constant for a specific reaction depends on many factors, such as temperature and pressure (for a gas).

A. Consider the following reaction: $CO_2 + H_2(g) \rightleftarrows CO(g) + H_2O(g)$

B. The equilibrium expression would be: $Kc = \dfrac{[CO]\,[H_2O]}{[CO_2]\,[H_2]}$

C. It is determined that, at equilibrium under specific conditions, the concentrations are:
$[CO_2] = 0.0954$ M $[H_2] = 0.0454$ M $[CO] = 0.0046$ M $[H_2O] = 0.0046$ M

D. Solve for Kc: $Kp = \dfrac{[0.0046]\,[0.0046]}{[0.0954]\,[0.0454\,]} = 4.9 \times 10^{-3}$

N.B. *Equilibrium constants do not have units.*

Model 2: Calculating Equilibrium Concentrations

Consider the following reaction: $H_2(g) + F_2(g) \rightleftarrows 2HF(g)$, given K = 64.0

A. The equilibrium expression would be: $K = 64.0 = \frac{[HF]^2}{[H_2]\,[F_2]}$

B. The initial concentrations are: $[H_2]$ = 2.00 M $[F_2]$ = 2.00 M [HF] = 0.00 M

C. Use the RICE (often called the ICE) table to solve for the concentrations at equilibrium:

Reaction	$[H_2]$	$[F_2]$	[HF]
Initial	2.00	2.00	0.00
Change	- x	- x	+ 2x
Equilibrium	2.00 – x	2.00 – x	0.00 + 2x

D. Solve: $K = 64.0 = \frac{[HF]^2}{[H_2]\,[F_2]} = \frac{[2x]^2}{[2.00-x]\,[2.00-x]}$

$$K = 64.0 = \frac{[2x]^2}{[2.00-x]^2}$$

$$8.00 = \frac{[2x]}{[2.00-x]}$$

(8.00)(2.00 – x) = 2x

(16.0) – (8.00x) = 2x

16.0 = 2x + 8.00x

1.60 = x

$[H_2] = [F_2]$ = 2.00 – 1.60 = 0.400 M

[HF] = 2(1.60) = 3.20 M

E. Check: $K = \frac{[HF]^2}{[H_2]\,[F_2]} = \frac{(3.20)^2}{(0.400)(0.400)} = 64$

Problems

1. At equilibrium: $[H_2] = [Cl_2]$ = 0.075 M, and [HCl] = 0.95 M. What is the equilibrium constant?
   $H_2(g) + Cl_2(g) \rightleftarrows \underline{2}HCl(g)$
2. At equilibrium: $[N_2] = [H_2]$ = 0.13 M, and [NO] = 0.19 M. What is the equilibrium constant?
   $N_2(g) + O_2(g) \rightleftarrows \underline{2}NO(g)$
3. A mixture of 0.750 M H_2 and 0.750 M I_2 are placed in a 1.00-L stainless-steel flask at 430°C. The equilibrium constant is 54.3. There is no HI at the beginning of the reaction. Calculate $[H_2]$, $[I_2]$, and [HI] at equilibrium. (Hint: rather than use the quadratic equation, take the square roots of both sides.)
   $H_2(g) + I_2(g) \rightleftarrows \underline{2}HI(g)$
4. Consider the reaction in question 3, but with the initial concentrations of $[H_2] = [I_2]$ = 0.040 M. Calculate $[H_2]$, $[I_2]$, and [HI] at equilibrium.
5. The initial concentrations are $[SO_2] = [NO_2] = [NO] = [SO_3]$ = 0.025 M. If Kc = 85.0, what are the concentrations at equilibrium? $SO_2(g) + NO_2(g) \rightleftarrows NO(g) + SO_3(g)$
6. For the reaction: $2NO(g) \rightleftarrows N_2(g) + O_{2(g)}$. At 300 K, the equilibrium constant is 2.43×10^3. Initially, 0.185 moles of each N_2 and O_2 are introduced into a 6.00 L container and allowed to reach equilibrium. What are the concentrations of N_2, O_2, and NO at equilibrium?
7. For the reaction, $HSO_4^- \rightleftarrows H^+ + SO_4^{2-}$; K = 0.012. Initially, only HSO_4^-, 0.500 M, is present. What are the concentrations of all three chemicals at equilibrium?