Chemical Equilibrium(≡m):

Consider the following reaction,

$$N_2(g) \quad + \quad 3H_2(g) \quad \leftrightarrow \quad 2NH_3(g)$$

Typical equilibriums,

$$N_2(g) \quad + \quad 3H_2(g) \quad \rightleftharpoons \quad 2NH_3(g)$$

$\rightleftharpoons$ denotes a reversible reaction and an equilibrium.

Ex:2

$$H_2O(l) \quad \rightleftharpoons \quad H_2O(g)$$

Equilibrium Constant(K):

$$K = \frac{[\text{products}]}{[\text{reactants}]}$$

[reactants] = concentration of reactants

[products] = concentration of products

For a general equilibrium:

$aA + bB \rightleftharpoons cC + dD$

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

For the equilibrium:

$4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g)$

$$K = \frac{[H_2O]^2[Cl_2]^2}{[HCl]^4[O_2]}$$

Equilibrium Constant(K_c):

K_c: Equilibrium constant where the concentration of substances are in molarity.

Ex: Calculate K_c for the following equilibrium

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

if at equilibrium(≡m) $[N_2O_4] = 4.27\times10^{-2}$ M and $[NO_2] = 1.41\times10^{-2}$ M.

Equilibrium Constant(K_c) cont…:

The magnitude of K_c indicates the position of the equilibrium.

$$K = \frac{[\text{products}]}{[\text{reactants}]}$$

$K_c > 1$ equilibrium favors formation of products.

$K_c < 1$ equilibrium favors formation of reactants.

Predicting the Direction of a Reaction:

The reaction quotient(Q) is used to determine if a system is at equilibrium.

$$Q = \frac{[\text{products}]}{[\text{reactants}]}$$

Q: Reaction quotient

Q is calculated the same as K_c.

Predicting the Direction of a Reaction cont…:

Ex: A ⇋ B $K_c = 5.0$

If [A] = 0.50 M and [B] = 4.0 M, calculate Q and predict the direction.

$Q = K_c$ Reaction at ≡m.

$Q > K_c$ Reaction will proceed from right to left to obtain ≡m.

$Q < K_c$ Reaction will proceed from left to right to obtain ≡m.

Example:

Ex:

If 0.100 mol $PCl_5(g)$, 0.0500 mol $PCl_3(g)$, and 0.0300 mol of Cl_2 is placed in a 1.00 L vessel, is the system at equilibrium, and if not which direction will the reaction move.

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) \qquad K_c = 0.0415$$

Homogeneous and Heterogeneous Equilibria

Homogeneous Equilibria- Equilibria between substances in the same phase.

Heterogeneous Equilibria- Equilibria between substances in two or more phases.

Examples:

Ex: $H_2O(l) \rightleftharpoons H_2O(g)$

$$K_c = [H_2O(g)]$$

Ex:2 $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

$$K_c = [CO_2(g)]$$

Ex:3

$3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$

$$K_c = \frac{[H_2(g)]^4}{[H_2O(g)]^4}$$

Calculating Equilibrium Concentrations:

K_c expressions can be used to determine equilibrium concentrations.

Ex:

$$H_2(g) \; + \; I_2(g) \rightleftharpoons 2HI(g) \qquad K_c = 54.5$$

A given amount of HI is placed in a 1.00 L vessel at 425°C. If the equilibrium concentration of HI is 0.50 M, calculate the equilibrium concentrations of H_2 and I_2.

Example:2

$$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g) \quad K_c = 0.771$$

If 0.0100 mol of H_2 and 0.0100 mol of CO_2 are mixed in a one litre vessel at 750 °C, what are the equilibrium concentration of all species at ≡m?

The Equilibrium Constant(K_p):

For the equilibrium

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

$$K_c = [CO_2(g)]$$

$$K_p = p_{CO_2}$$

Relationship between K_p and K_c:

$$K_p = K_c (RT)^{\Delta n}$$

Δn = moles of gas products – moles of gas reactants

T = temperature in Kelvins

R = Gas Constant(0.0821 L·atm/K·mole)

Ex:

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

$\Delta n = 2 - 1 = +1$

$$K_p = K_c (RT)^1$$

$$K_p = K_c RT$$

Ex:2

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$$

Δn = 1 – 2 = -1

$$K_p = K_c(RT)^{-1}$$

$$K_p = \frac{K_c}{RT}$$

Ex:3

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

$\Delta n = 2 - 2 = 0$

$$K_p = K_c (RT)^0$$

$$K_p = K_c$$

Le Chatelier’s Principle:

When an equilibrium is disturbed it will cause the equilibrium to shift in a direction that minimizes the effect and establish a new equilibrium.

Concentration Changes

Pressure Changes

Temperature Changes