

Chemical Equilibrium(≡m):

Consider the following reaction,



Typical equilibriums,



\rightleftharpoons denotes a reversible reaction and an equilibrium.

Ex:2



Equilibrium Constant(K):

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

[reactants] = concentration of reactants

[products] = concentration of products

For a general equilibrium:



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

For the equilibrium:



$$K = \frac{[\text{H}_2\text{O}]^2 [\text{Cl}_2]^2}{[\text{HCl}]^4 [\text{O}_2]}$$

Equilibrium Constant(K_c):

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

Ex: Calculate K_c for the following equilibrium



if at equilibrium($\equiv m$) $[\text{N}_2\text{O}_4] = 4.27 \times 10^{-2} \text{ M}$
and $[\text{NO}_2] = 1.41 \times 10^{-2} \text{ M}$.

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...:



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

$Q = K_c$ Reaction at $\equiv m$.

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

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

Example:

Ex:

If 0.100 mol $\text{PCl}_5(\text{g})$, 0.0500 mol $\text{PCl}_3(\text{g})$, 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.



Homogeneous and Heterogeneous Equilibria

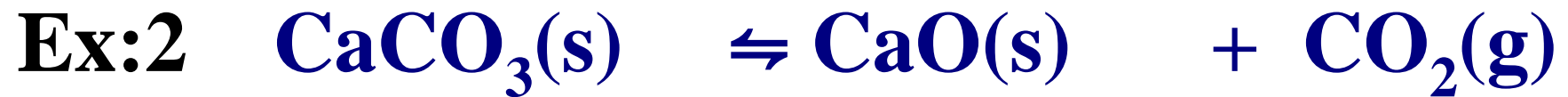
Homogeneous Equilibria- Equilibria between substances in the same phase.

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

Examples:

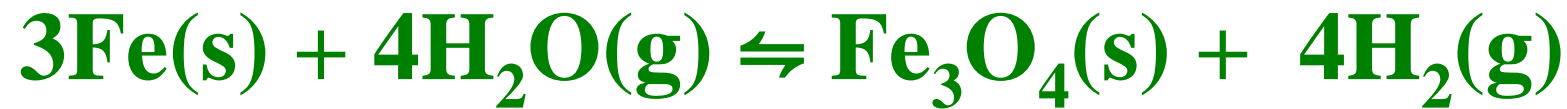


$$K_c = [\text{H}_2\text{O}(\text{g})]$$



$$K_c = [\text{CO}_2(\text{g})]$$

Ex:3

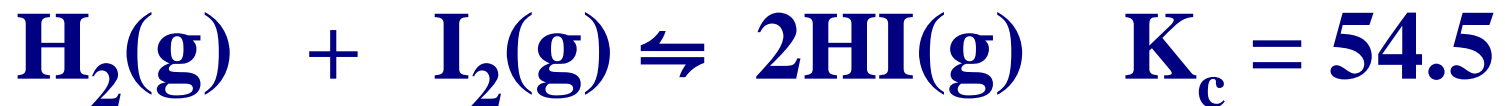


$$K_c = \frac{[\text{H}_2(\text{g})]^4}{[\text{H}_2\text{O}(\text{g})]^4}$$

Calculating Equilibrium Concentrations:

K_c expressions can be used to determine equilibrium concentrations.

Ex:



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 H₂ and I₂.

Example:2



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 $\equiv m$?

The Equilibrium Constant(K_p):

For the equilibrium



$$K_c = [\text{CO}_2(\text{g})]$$

$$K_p = p_{\text{CO}_2}$$

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:



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

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

$$K_p = K_c RT$$

Ex:2



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

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

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

Ex:3



$$\Delta n = 2 - 2 = 0$$

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

$$K_p = K_c$$

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