Equilibrium

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Reversible Reactions

In most chemical reactions, the chemical reaction can be reversed, where the products will form the reactants.

This type of reversible reaction is quite common, and occur sometimes spontaneously and sometimes require different conditions to occur.

Chemical Equilibrium

A chemist will use reversible reactions to demonstrate a concept of equilibrium.

After a period of time, a reaction's forward reaction will occur at the same rate as its reverse reaction. When this occurs, the concentration of the products does not increase or decrease, nor does the reactants.

In other words, chemical equilibrium occurs when the concentrations of the reactants and products remain constant because the rate of the forward reaction equals the rate of the reverse reaction.

Equilibrium Constant

Guldberg and Waage developed a theory to determine when equilibrium will occur. Their experiment measured the concentrations at equilibrium, and discovered that every reaction has its own equilibrium constant (Kc).

This equilibrium constant is always a constant value for a given reaction regardless of the initial concentration.

Writing the Kc

Based on the reaction:

$$aA + bB <==> cC + dD$$

the K_c is written as: $\frac{[C]^c[D]^d}{[A]^a[B]^b}$

In addition to the previous rules, only concentrations of gases and aqueous solutions will be used in the K_c

Reason: Concentrations of solids and liquids normally do not change that much in reversible reactions.

Equilibrium Constant

The equilibrium constant can also be listed in terms of pressure of gas of the materials. For the general reaction:

$$aA + bB <==> cC + dD$$

the K_p is written as: $\frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$

Where P_C is the partial pressure of gas C, etc. The relationship between K_p and K_c is:

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

Calculating K

When you know all the concentrations or pressures of a reaction at equilibrium, it is possible to calculate the equilibrium constant.

Example: For the reaction

$$H_2\ (g) + I_2\ (g) <==> 2\ HI\ (g)$$

the equilibrium concentrations are $[H_2]$ = 0.95 M, $[I_2]$ = 0.95 M and [HI] = 7.1 M at 20 ^{O}C. What is the equilibrium constant for the reaction?

Problems involving K

Occasionally, only one concentration is given at equilibrium. However, it is possible to determine the other concentrations in reaction to determine K.

If the K_p for the reaction N_2O_4 (g) <==> 2 NO_2 (g) is 4.73, and the equilibrium pressure of $[N_2O_4]$ is 0.45 atm, what is the equilibrium pressure of NO_2?

Problems involving K

The reaction 2 O_3 (g) <==> 3 O_2 (g) starts with a mixture of 0.50 mol of O_3 and 0.50 mol of O_2 in a 1.0 L flask. The reaction reacts, and at equilibrium, $[O_2]$ = 0.86 M.

a) What is the equilibrium concentration of O_3?

b) What is the K_c for the reaction?

Problems involving K

The reaction H_2 (g) + I_2 (g) <==> 2 HI (g) starts with a mixture of gases with concentrations for H_2 and I_2 as 2.0 M and 3.0 M, respectively. The K_c for the reaction is 50.5 .

What are the equilibrium concentrations for all three materials?

Reaction Quotient

When the reactants and products of a reaction are mixed, it is hard to tell if the system is at equilibrium.

From the mixing concentrations though, we can predict which direction the reaction will proceed.

We can calculate the reaction quotient (Q), which is calculated the same as the K_c, and will determine the direction of the reaction.

Reaction Quotient

If the K_c for the reaction H_2(g) + I_2(g)<==> 2 HI (g) is 51, in what direction will the reaction be heading when the amounts of each gas are found to be 2.0 x 10^{-2} mol HI, 1.0 x 10^{-2} mol H_2, and 3.0 x 10^{-2} mol I_2 in a 2.0 L flask?

Q and K_{eq}

1) If $Q = K_c$, then the reaction is at equilibrium.
2) If $Q > K_c$, then there are more products than reactants, and the reaction will shift to the left.
3) If $Q < K_c$, then there are more reactants than products, and the reaction will shift to the right.

Magnitude of K

When K is greater than 1, what does that tell us about the amounts of product and reactant at equilibrium?

What if K is less than 1?

Le Chatelier's Principle

If a change in conditions is imposed on a system at equilibrium, the equilibrium will shift in the direction that will reduce the imbalance.

1) Changed concentration

- Adding more of one substance will shift the equilibrium away from that substance.
- Taking out some of one substance will shift the eq. toward that substance.

More Le Chatelier

2) Changed Pressure

- Increasing pressure will shift it toward less moles of gas
- Decreasing pressure will shift it toward more moles of gas
- Will not affect aqueous equilibria

3) Changed temperature

- Increasing temperature will shift it toward a more endothermic reaction
- Decreasing temperature will shift it toward a more exothermic reaction