ANALYTICAL CHEMISTRY

Read the details of the information given below from Skoog and West's "Fundamentals of Analytical Chemistry" book, which is recommended as a reference.

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Aqueous Solutions and Chemical Equilibria

The Chemical Composition of Aqueous Solutions

Chemical Equilibrium

Buffer Solutions

9B Chemical equilibrium

Equilibrium-constant expressions are equations that describe the concentration relationships among reactants and products at equilibrium.

The equilibrium state

$$H_3AsO_4 + 3I^- + 2H^+ \rightleftharpoons H_3AsO_3 + I_3^- + H_2O$$

3 mmol I^- added to 1 mmol H_3AsO_4

The concentration relationship at chemical equilibrium is independent of the route to the equilibrium state.

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This relationship is altered by applying stress to the system.

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changes in temperature, in pressure or in total concentration of a reactant or a product

Le Châtelier's principle states that the position of an equilibrium always shifts in such a direction as to relieve a stress that is applied to the system.

Equilibrium-constant expressions

The influence of concentration or pressure (if the participants are gases) on the position of a chemical equilibrium is described in quantitative terms by means of an **equilibrium-constant expression**.

$$wW + xX \rightleftharpoons yY + zZ$$

$$K = \frac{[Y]^y[Z]^z}{[W]^w[X]^x}$$

The constant K is a temperature-dependent numerical quantity called the **equilibrium constant**.

Using solubility-product constants

When we say that a sparingly soluble salt is completely dissociated, *we do not imply* that all of the salt dissolves.

What we mean is that the very small amount that *does* go into solution dissociates completely.

$\boxed{K_{sp}\text{ : the }\textbf{solubility-product constant}}$

The solubility of a precipitate in pure water

With the solubility-product expression, solubility of a sparingly soluble substance that ionizes completely in water can be calculated.

The effect of a common ion on the solubility of a precipitate

The solubility of an ionic precipitate decreases when a soluble compound containing one of the ions of the precipitate is added to the solution.

This behavior is called the **common-ion effect**.

Buffer solutions

Buffers are used in all types of chemical applications whenever it is important to maintain the pH of a solution at a constant and predetermined level.

Calculating the pH of buffer solutions

$$HA + H_2O \rightleftharpoons H_3O^+ + A^- \quad K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$A^- + H_2O \rightleftharpoons OH^- + HA \quad K_b = \frac{[OH^-][HA]}{[A^-]} = \frac{K_w}{K_a}$$

$K_a > K_b$	the solution is acidic
$K_a < K_b$	the solution is basic

Buffer capacity

The **buffer capacity**, β, of a solution is defined as the number of moles of a strong acid or a strong base that causes 1.00 L of the buffer to undergo a 1.00-unit change in pH.