

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



3 mmol I⁻ added to 1 mmol H₃AsO₄

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



This relationship is altered by applying stress to the system.



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**.



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

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.

K_{sp} : the **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

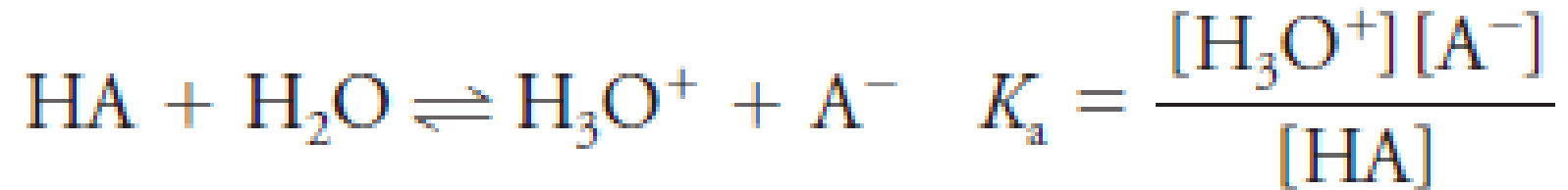
The solubility of an ionic precipitate decreases when a soluble compound containing one of the ions of the precipitate is added to the solution.

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



$$K_a > K_b$$

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.