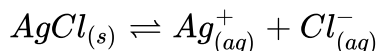


Solubility and Acid-Base

Solubility



The **solubility product** (K_{sp}) of this reaction would be $K_{sp} = [Ag^{+}][Cl^{-}]$

We can use ICE tables (see previous outline) to solve solubility problems.

To predict whether or not a reaction will precipitate, we can calculate the Q_{sp} and compare to the K_{sp} .

$$\begin{aligned} Q_{sp} &= K_{sp} && \text{Equilibrium} \\ Q_{sp} &< K_{sp} && \text{No precipitation} \\ Q_{sp} &> K_{sp} && \text{Precipitation occurs} \end{aligned}$$

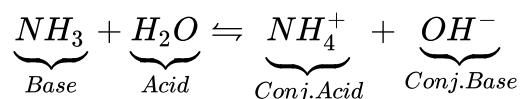
when $Q_{sp} > K_{sp}$ The system becomes super saturated and as a result, precipitate forms.

The **common ion effect** is simply the application of Le Chatliers principle to a solution that already contain an ion involved in a dissolution reaction. e.g. if we were to predict the shift in the above equation in a sodium chloride solution rather than pure water, the equilibrium would shift to precipitate because there is already chloride present, making the AgCl less soluble

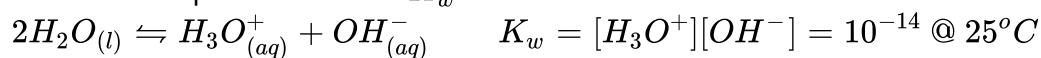
Brønsted-Lowry acids and bases

In acid-base theory, a **Brønsted-Lowry acid** is a molecule that gives up a proton (H^{+}) and a **Brønsted-Lowry base** is a molecule that accepts a proton. A **conjugate base** is the basic anion that occurs upon deprotonating an acid and a **conjugate acid** is the cation that occurs when a base accepts a proton.

Example:



Water is an example of an **amphoteric** substance, it is capable of either donating or accepting a proton. It can also **autoionize** in which 2 water molecules can form a charged hydronium ion (H_3O^{+}) and hydroxide ion (OH^{-}). The reaction and ion product constant K_w are shown below.



pH and pOH

pH and **pOH** are measures of the acidity and basicity of a solution respectively. Both are simply the negative logarithm base 10 of either the hydronium or hydroxide ion concentration in the solution.

$$pH = -\log([H_3O^{+}])$$

The pH scale is a common metric of gauging the acidity of solutions. Values less than 7 are acidic, values higher than 7 are basic, with 7 being the pH of pure water. A similar metric, pOH exists for hydroxide;

$$pOH = -\log([OH^{-}])$$

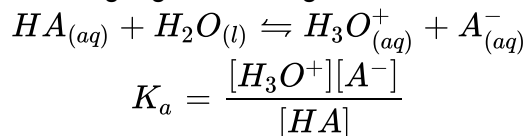
If we take the negative log of K_w we find;

$$K_w = [H_3O^{+}][OH^{-}]$$

$$\begin{aligned}
 -\log(K_w) &= -\log([H_3O^+]) + (-\log([OH^-])) \\
 pK_w &= pH + pOH \\
 14 &= pH + pOH
 \end{aligned}$$

Relative Strengths of Acids and Bases

K_a is the acid-ionization constant and can gauge how strong an acid is. For the below reaction



The base-ionization constant, K_b , is the same, but for bases.

Relating these values to K_w for conjugate pairs;

$$K_a \times K_b = \frac{[H_3O^+][A^-]}{[HA]} \times \frac{[HA][OH^-]}{[A^-]} = [[H_3O^+][OH^-]] = K_w$$

Since a strong acid will completely dissociate, $[HA]$ will be 0, and thus K_a will approach infinity. If $K_a = \text{infinity}$, then if $K_b = K_w/K_a = K_w/\infty = 0$

Polyprotic Acids

Acids that can donate one proton are called **monoprotic**, acids that can donate multiple protons are **polyprotic** (di,tri,etc). For a polyprotic acid, each proton has it's own associated K_a value (K_{a1} , K_{a2} , etc.)

Buffers

Buffers are mixtures of roughly the same amount of a weak acid/base and it's conjugate (acid/conjugate base or vis a versa) such that the solution resists pH changes from adding small amounts of acids or bases to the solution. If acid is added to an acid/conjugate base buffer, then the equilibrium would shift left, away from acidity, if base is added, the equilibrium shifts right towards acidity, effectively neutralizing a pH change.

Henderson-Hasselbeck Equation

$$\begin{aligned}
 K_a &= \frac{[H_3O^+][A^-]}{[HA]} \\
 [H_3O^+] &= K_a \times \frac{[HA]}{[A^-]}
 \end{aligned}$$

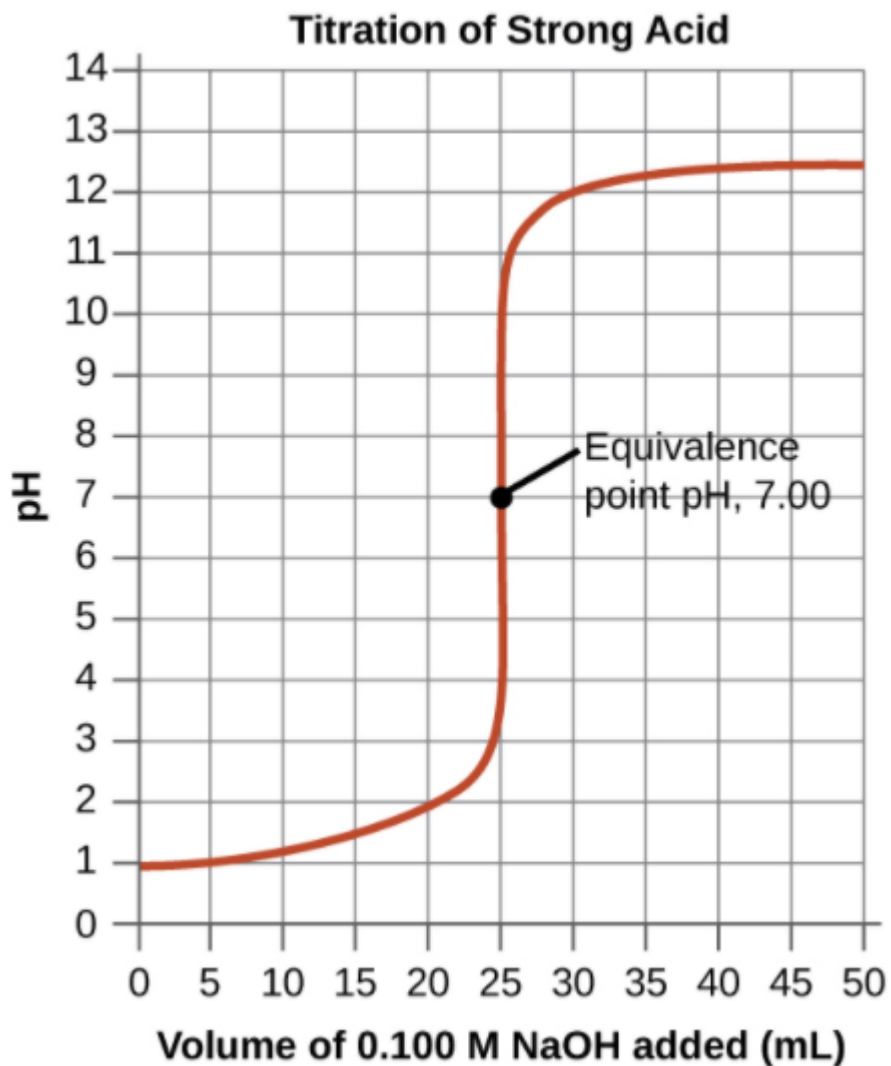
take the $-\log()$ of the whole thing;

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

Titration

A **titration** curve is a plot of pH as a function of addition of a **titrant**.

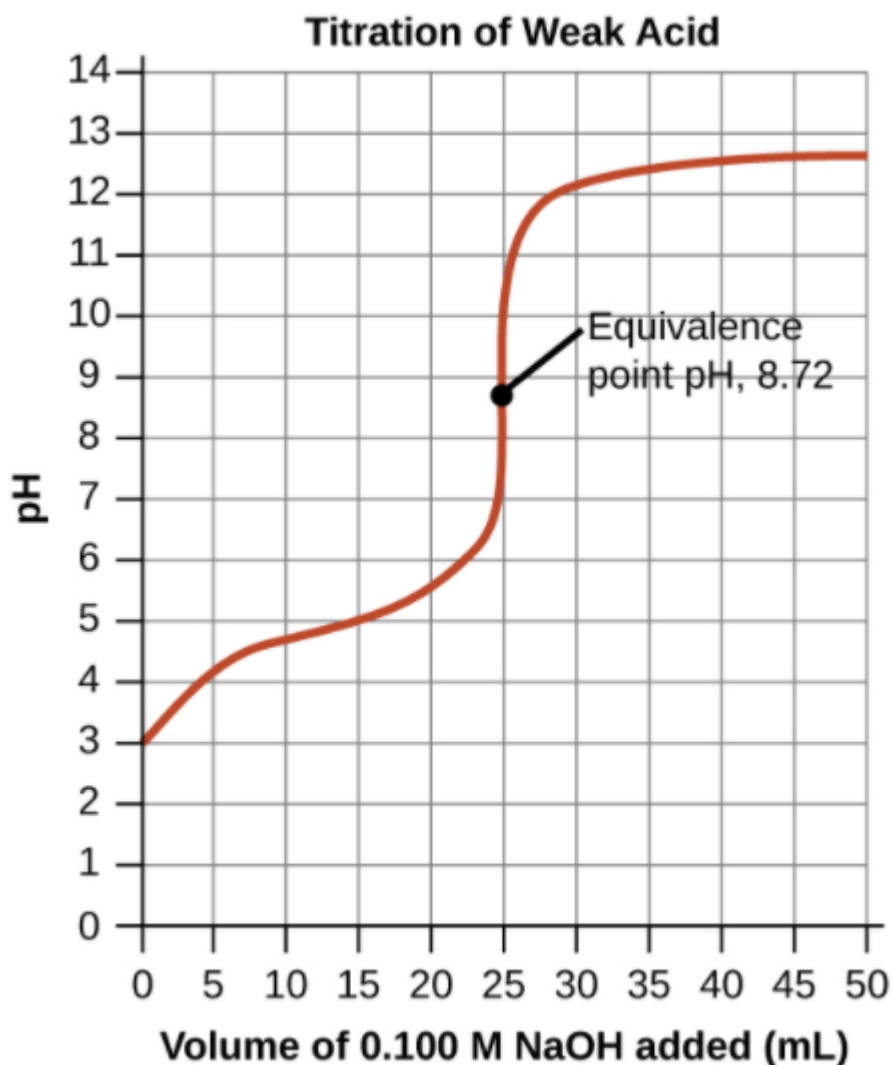
Strong Acid/Strong Base



Take an example of a 25 mL sample of 0.1 M HCl being titrated by 0.1 M NaOH. They both fully dissociate and the molar ratio of the 2 molecules in the neutralization reaction is 1 to 1 (i.e. 1 proton to 1 hydroxide). Since this is the case, once you add 25 mL of base, the solution will reach the equivalence point at a pH of 7.

Weak Acid/Strong Base

Example 25 mL of 0.1 CH₃COOH and 0.1 mL NaOH.



Since the acid dissociates based on its K_a value (1.8×10^{-5}), we can mathematically calculate the equivalence point using an ICE method.

We start with $0.025\text{ L} \times 0.1\text{ M CH}_3\text{COOH} = 0.0025\text{ mol CH}_3\text{COOH}$

After adding 25 mL of NaOH, we have $[\text{CH}_3\text{COOH}] = 0.0025\text{ mol} / 0.05\text{ L} = 0.05\text{ M}$

$$K_b = K_w / K_a = \frac{10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

Using ICE, assuming that the dissociation is sufficiently small;

$$K_b = \frac{x^2}{0.05}$$

$$x = 5.3 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pOH} = -\log(5.3 \times 10^{-6}) = 5.28$$

$$\text{pH} = 14 - \text{pOH} = 14 - 5.28 = 8.72$$

At half the equivalence point, half of the acid has been neutralized by the base, so that the concentrations of conjugate base is equal to that of the acid;

$$pH = pK_a + \log \frac{[CH_3COO^-]}{[CH_3COOH]} = pK_a + \log(1) = pK_a$$

Acid-Base Indicators

Acid-base indicators are weak organic acids or bases that experience a color change at a specific pH. Typically one will select an indicator that has a color change in the range of the region of interest.

