

Section 14.1 The Nature of Acids and Bases

Arrhenius concept:

Acid – produces hydrogen ions in aqueous solutions.

Bases – produces hydroxide ions in aqueous solutions.

Bronsted – Lowery model

Acid – a proton (H⁺) donor.

Base – a proton acceptor

Hydronium ion the ion that forms when water accepts a proton (H⁺)

General reaction that occurs when an acid is dissolved in water.



Conjugate Base everything that remains of acid molecule after a proton is lost.

Conjugate Acid the molecule which is formed when the proton is transferred to the base.

Conjugate Acid - Base pair consists of two substances related to each other by the donating and accepting of a single proton.

Equilibrium expression for a general acid dissolved in water reaction would be

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

K_a is called the **Acid Dissociation Constant**.

Section 14.2 Acid Strength

Strong Acid - one in which the equilibrium lies far to the right. It has a large value for K_a . Strong acids also yield weak conjugate bases.

Weak Acid - one in which the equilibrium lies far to the left. It has a small value for K_a . Weak acids yield relatively strong conjugate bases. See Table 14.1 p.651

Diprotic Acid - an acid having two acidic protons. ex. - H₂SO₄

Oxyacids - the acidic proton is attached to an oxygen atom.

Organic Acids - those acids with a carbon atom backbone, and commonly contain the carboxyl group.

Monoprotic Acid - acids having only one acidic proton.

Amphoteric - can behave either as an acid or as a base. ex. - water

Autoionization - involves the transfer of a proton from one water molecule to another to produce a hydroxide ion and a hydronium ion.



**Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.

Ion-Product Constant K_w (or dissociation constant for water) - always refers to the autoionization of water. Using the above reaction for the autoionization of water:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = [\text{H}^+][\text{OH}^-]$$

at 25°C

$$[\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}$$

so at 25°C

$$K_w = (1.0 \times 10^{-7} \text{ M})^2 = 1.0 \times 10^{-14} \text{ mol}^2/\text{L}^2$$

It is important to recognize the meaning of K_w . The product of $[\text{H}^+]$ and $[\text{OH}^-]$ must always equal 1.0×10^{-14} . This lends itself to three possibilities:

1. A neutral solution, where $[\text{H}^+] = [\text{OH}^-]$.
2. An acidic solution, where $[\text{H}^+] > [\text{OH}^-]$.
3. A basic solution, where $[\text{H}^+] < [\text{OH}^-]$.

Section 14.3 The pH Scale

pH Scale - provides a convenient way to represent solution acidity

$$\text{pH} = -\log [\text{H}^+]$$

Significant Figures for Logarithms - the number of decimal places in the log is equal to the number of significant figures in the original number.

Similar scales can be formed for pK and pOH.

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

$$\text{pK} = -\log K$$

Many relationships may be formed between these three equations - useful equations to know

$$\begin{aligned} K_w &= [\text{H}^+][\text{OH}^-] \\ \log K_w &= \log [\text{H}^+] + \log [\text{OH}^-] \\ \text{p}K_w &= \text{pH} + \text{pOH} \end{aligned}$$

and at 25°C

$$\text{pH} + \text{pOH} = 14.00$$

Section 14.4 Calculating the pH of Strong Acid Solutions

Common Strong Acids - (commit to memory) - $\text{HCl}(\text{aq})$, $\text{HNO}_3(\text{aq})$, $\text{H}_2\text{SO}_4(\text{aq})$, and $\text{HClO}_4(\text{aq})$

One key to solving acid-base equilibrium, is you must focus on the solution components and their chemistry.

Major Species - are those solution components present in relatively large amounts.

Section 14.5 Calculating the pH of Weak Acid Solution

Step 1 - List the major species in the solution.

Step 2 - Choose the species that can produce H^+ , and write a balanced equation for the reaction producing H^+ .

Step 3 - Using the values of the equilibrium constant for the reactions you have written, decide which equilibrium will dominate in producing H^+ .

Step 4 - Write the equilibrium expression for the dominant equilibrium.

**Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.

- Step 5 - List the initial concentrations of the species participating in the dominant equilibrium.
 Step 6 - Define the change needed to achieve equilibrium; that is, define x.
 Step 7 - Write the equilibrium concentrations in terms of x.
 Step 8 - Substitute the equilibrium concentrations into the equilibrium expression.
 Step 9 - Solve for x the "easy" way: that is by assuming $[HA]_0 - x \approx [HA]_0$
 Step 10 - Using the 5% rule verify whether the approximation is valid.
 Step 11 - Calculate $[H^+]$ and pH

5% Rule - $\frac{x}{[HA]_0} \times 100 < 5\%$ then the approximation is acceptable.

Percent Dissociation - $\text{Percent dissociation} = \frac{\text{amount dissociated} \left(\frac{\text{mol}}{\text{L}}\right)}{\text{initial concentration} \left(\frac{\text{mol}}{\text{L}}\right)} \times 100\%$

For a given weak acid, the percent dissociation increases as the acid becomes more dilute. For solutions of any weak acid HA, $[H^+]$ decreases as $[HA]_0$ decreases, but the percent dissociation increases as $[HA]_0$ decreases.

Section 14.6 Bases

Strong Base - one which dissociates completely in water and has large K_b values.

General reaction for a base B and water is given by:



The equilibrium constant for this equation:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

Weak Base - small K_b values.

Section 14.7 Polyprotic Acids

Polyprotic Acids - acids which can furnish more than one proton. Table 14.4 p.677 - important K_a values.

Characteristics of Weak Polyprotic Acids

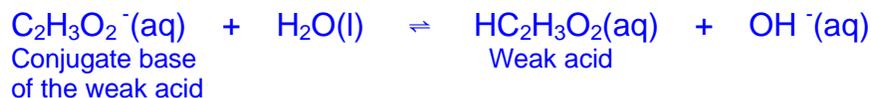
- Typically, successive K_a values are so much smaller than the first value that only the first dissociation step makes a significant contribution to the equilibrium concentration of H^+ . This means that the calculation of the pH for a solution of a typical weak polyprotic acid is identical to that for a solution of a weak monoprotic acid.
- Sulfuric acid is unique in being a strong acid in its first dissociation step and a weak acid in its second step. For relatively concentrated solutions of sulfuric acid (1.0 M or higher), the large concentration of H^+ from the first dissociation step represses the second step. Which can be neglected as a contributor of H^+ ions. For dilute solutions of sulfuric acid, the second step does make a significant contribution, and the quadratic equation must be used to obtain the total H^+ concentration.

Section 14.8 Acid - Base Properties of Salts

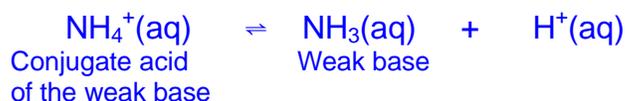
Salt - another name for an ionic compound.

Salts that produce neutral solutions - salts that consist of the cations of strong bases and the anions of strong acids have no effect on $[H^+]$ when dissolved in water.

Salts that produce basic solutions - any salt whose cation has neutral properties and whose anion is the conjugate base of a weak acid, the aqueous solution will be basic. Ex:



Salts that produce acidic solutions - salts in which the anion is not a base and the cation is the conjugate acid of a weak base produce acidic solutions. Ex:



Highly charged metal ions - salts which contain highly charged metal ions produce an acidic solution. The metal ion becomes hydrated which then causes the solution to become acidic. See p. 687 - 689 (and Table 14.6)

Section 14.9 The Effect of Structure on Acid-Base Properties

There are two main factors that determine whether a molecule containing an X - H bond will behave as a Bronsted-Lowry acid: the strength of the bond and the polarity of the bond. Increased polarity and high electron density typically leads to large K_a values (strong acids).

Section 14.10 Acid-Base Properties of Oxides

A compound containing the H - O - X group will produce an acidic solution in water if the O - X bond is strong and covalent. If the O - X bond is ionic, the compound will produce a basic solution in water.

Acidic Oxides - a covalent oxide dissolves in water, and an acidic solution forms.

Basic Oxides - an ionic oxide dissolves in water, and oxide has a great affinity for H^+ causing basic solutions.

Section 14.11 The Lewis Acid-Base Model

Lewis acid - is an electron-pair acceptor.

Lewis base - is an electron-pair donor.

The Lewis model encompasses the Bronsted-Lowry model, but the reverse is not true.

Section 14.12 Strategy for Solving Acid-Base Problems: A Summary

Read through pages 695 - 696

**Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.