10.1 Acids and Bases in Aqueous Solution

Arrhenius Definition of Acids and Bases

- An acid is a substance that gives hydrogen ions, H⁺, when dissolved in water. In fact, H⁺ reacts with water and produces H₃O⁺.
- A base is a substance that gives hydroxide ions, OH⁻, when dissolved in water.
- The neutralization reaction of an acid with a base yields water plus a salt.

10.2 Some Common Acids and Bases

Some common acids are: Sulfuric acid, H₂SO₄; Hydrochloric acid, HCl Phosphoric acid, H_3PO_4 ; Nitric acid, HNO₃ Acetic acid, CH₃CO₂H Some common bases are: Sodium hydroxide, NaOH; Calcium hydroxide, Ca(OH)₂ Magnesium hydroxide, $Mg(OH)_2$ Ammonia, NH₃

The Bronsted-Lowry Definition of Acids and Bases *Bronsted-Lowry Acids*: Any substance that is able to give a hydrogen ion, H⁺, to another ion or molecule. H⁺ ions are also known as a *proton*. Therefore, acids are those substances that can donate protons.



Monoprotic acid : Acids with one proton to donate, such as, hydrochloric acid, HCl; nitric acid, HNO_3

Diprotic acid: Acids with two protons to donate, such as sulfuric acid, H_2SO_{4} .

Triprotic acid: Acids with three protons to donate, such as phosphoric acid, H_3PO_4 .

Bronsted-Lowry bases: Any substance that is able to accept a hydrogen ion, H⁺, from an acid. A base can be neutral or negatively charged, for example, ammonia, NH_3 and hydroxide ion, OH^- . Products of an acid-base reactions are themselves acids and bases.

Conjugate acid-base pair: Two substances whose formula differ by only a hydrogen ion, H⁺.

Conjugate base: The substance formed by loss of H⁺ from an acid.

Conjugated base: The substance formed by addition of H⁺ to a base.

10.4 Water as Both an Acid and a Base

Water is neither an acid nor a base according to Arrhenius acid-base theory since water does not contain appreciable amount of H_3O^+ or OH^- . However, according to Bronsted-Lowry acid-base theory, water is both an acid and a base.



10.5 Some Common Acid-Base Reactions

Reaction of Acids with Hydroxide Ion:

Acids react with metal hydroxides to produce water and a salt in a neutralization reaction.

Reactions of Acids with Bicarbonate and carbonate Ion

Bicarbonate ion, HCO_3^- , reacts with acid by accepting H⁺ to yield carbonic acid, H_2CO_3 . Similarly, carbonate ion accepts two protons in its reaction with acid. H_2CO_3 reaction product is unstable, rapidly decomposing to yield carbon dioxide gas and water.

Reaction of Acids with Ammonia

Acids react with ammonia to yield ammonium salts, such as ammonium chloride. Most of the ammonium salts are water soluble.

Living organisms contain a group of compounds called amines, which contain ammonia-like nitrogen atoms bonded to carbon. Amines react with acids just as ammonia does, yielding water soluble salts.

10.6 Acid and Base Strength

Acids differ in their ability to give up a proton. Similarly, bases differ in their ability to accept a proton. Acids and bases are classified as follows:

- Strong acid: An acid that gives up H⁺ easily and is essentially 100% dissociated (splits to produce H⁺ and an anion) in water.
- Weak acid: An acid that gives up H⁺ with difficulty and is less than 100% dissociated in water.

- Strong base: A base that has a high affinity for H⁺ and holds it tightly.
- Weak base: A base that has only a slight affinity for H⁺ and holds it weakly.

The stronger the acid, the weaker its conjugate base; the weaker the acid, the stronger its conjugate base.



	Acid			Conjugate Base			
Increasing acid strength	Strong acids: 100% dissociated	Perchloric acid Sulfuric acid Hydriodic acid Hydrobromic acid Hydrochloric acid	HClO ₄ H ₂ SO ₄ HI HBr HCl HNO	ClO_4^- HSO_4^- I^- Br ⁻ Cl ⁻	Perchlorate ion Hydrogen sulfate ion Iodide ion Bromide ion Chloride ion	Little or no reaction as bases	Increasing base strength
		Hydronium ion	нцо ₃ Н ₃ O+	H ₂ O	Water		
	Weak acids	Hydrogen sulfate ion Phosphoric acid Nitrous acid Hydrofluoric acid	HSO ₄ - H ₃ PO ₄ HNO ₂ HF	SO ₄ ²⁻ H ₂ PO ₄ ⁻ NO ₂ ⁻ F ⁻	Sulfate ion Dihydrogen phosphate ion Nitrite ion Fluoride ion	Very weak bases	
	Very weak acids	Carbonic acid Dihydrogen phosphate ion Ammonium ion Hydrocyanic acid Bicarbonate ion Hydrogen phosphate ion	$H_{2}CO_{3}$ $H_{2}PO_{4}^{-}$ NH_{4}^{+} HCN HCO_{3}^{-} HPO_{4}^{2-}	HCO_{3}^{-} HPO_{4}^{2-} NH_{3} CN^{-} CO_{3}^{2-} PO_{4}^{3-}	Bicarbonate ion Hydrogen phosphate ion Ammonia Cyanide ion Carbonate ion Phosphate ion	Weak bases	
		Water	H ₂ O	он-	Hydroxide ion	Strong base	

TABLE 10.1 Relative Strengths of Acids and Conjugate Bases

Table 10.1 Relative strengths of acids and bases

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Chapter Ten

10.7 Acid Dissociated Constants

The reaction of weak acid with water can be described by an equilibrium equation. Equilibrium constant, K, and water concentration $[H_3O^+]$ together makes the acid dissociation constant k_a . Acid dissociation constant is a measure of acid strength.

- Strong acids have k_a value much greater than
 1.
- Weak acids have k_a value much less than 1.
- Donation of each successive H⁺ from a polyprotic acid is more difficult than the one before it, so ka value become successively smaller.
- Most organic acids, which contains $-CO_2H$ group, have k_a value near 10⁻⁵.

10.8 Dissociation of Water

Like all weak acids, water is slightly dissociated into H^+ and OH^- ions. The concentrations of the two ions are identical. At 25°C, concentration of each ion is 1.00×10^{-7} .

Ion product constant for water, k_w:

$$k_w = k_a[H_2O] = [H_3O^+][OH^-]$$

= [1.00 x 10⁻⁷][1.00 x 10⁻⁷]
= 1.00 x 10⁻¹⁴ at 25°C.

Product of $[H_3O^+]$ and $[OH^-]$ is a constant. Therefore, in an acidic solution where $[H_3O^+]$ is large and $[OH_-]$ must be small.

10.9 Measuring Acidity in Aqueous Solution: pH

A pH value between 0 and 14 is used to indicate concentration of H_3O^+ or OH^- in solution. Mathematically, the pH of a solution is defined as the negative common logarithm of the H_3O^+ concentration:

 $pH = -log [H_3O^+] or$ $[H_3O^+] = 10^{-pH}$

Acidic solution: pH < 7 $[H_3O^+] > 1.00 \ge 10^{-7} M$ Neutral solution: pH = 7 $[H_3O^+] = 1.00 \ge 10^{-7} M$ Basic solution: pH > 7 $[H_3O^+] < 1.00 \ge 10^{-7} M$



Fig 10.2 The pH scale and pH ofFig 10.3 relationship of pH tosome common substances H^+ and OH^- ion concentrations

10.10 Working with pH

- Calculate [H₃O⁺] in blood that has a pH of 7.4.
 -log [H₃O⁺] = pH; rearrangement of this equation gives [H₃O⁺] = antilog pH
 Therefore, [H₃O⁺] = antilog (-7.4) = 4 x 10⁻⁸
- Calculate pH of a solution that has [H₃O⁺] of 4.6 x 10⁻³
 -log [H₃O⁺] = pH
 -log (4.6 x 10⁻³) = - (-2.34) = 2.34

10.11 Laboratory Determination of Acidity

- The simplest but least accurate way to measure the pH of a solution is to use an acid-base indicator that changes color depending on the pH of the solution. For example:
- *Litmus* is red below pH 4.8 but blue above pH 7.8.
 - *Phenolphthalein* is colorless below pH 8.2 but red above pH 10;
- A mixture of indicators known as *universal indicator* gives approximate pH measurements in the range 2-10 pH.
- For more accurate pH measurements, an electronic pH meter can be used.

10.12 Buffer Solutions

Buffer: A combination of substances that act together to prevent a drastic change in pH. Usually mixture of a weak acid and its conjugate base is used to prepare a buffer solution.

For example, a solution that contains 0.10 M acetic acid and 0.10 M acetate ion can be utilized as a buffer, Fig 10.6.

If a small amount of OH⁻ is added to the above buffer solution, the pH will increase, but not very much because acetic acid of the buffer mixture will neutralize the added OH⁻. If small amount of H_3O^+ is added to the buffer, the pH will decrease, but again not very much since acetate ion of the buffer mixture will neutralize the added H_3O^+ .



Fig 10.6 pH changes in pure water and a buffer solution

Buffers should meet the following requirements to be the most effective:

- The ration of [HA] to [A⁻] should be close to
 1.
- The amounts of [HA] and [A⁻] in the buffer should be approximately 10 times greater than either the amount of acid or base you expect to add.

10.13 Buffers in the Body

The pH of body fluid is maintained by the following three major buffer systems:

- Carbonic acid bicarbonate system
- Dihydrogen phosphate hydrogen phosphate system
- Proteins acting as either proton acceptor or proton donors at different pH values.

Relationships of the bicarbonate buffer system to the lungs and kidneys shown in the following figure 10.7.



Fig 10.7 Lungs and kidneys relation with the bicarbonate buffer system

10.14 Acid and Base Equivalents

One equivalent of acid is equal to the molar mass of the acid divided by the number of H^+ ion produced per formula unit. Similarly, one equivalent of a base is the weight in grams that can produced one mole of OH^- ions. For example;

One equivalent of the monoprotic acid HCl is 36.5 g, which comes from the molar mass of the acid (36.5 g) divided by 1 (the number of H⁺ ion HCl can produced).

One equivalent of diprotic acid H_2SO_4 is 49.9 g. Molar mass of H_2SO_4 (98.0 g) divided by 2 (number of H⁺ sulfuric acid can produced)

