University of Al-Qadisiyah

College of Dentistry

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# MEDICAL CHEMISTRY

## Lecturer 1

## Acid, Base, and Salt

## Introduction and General Concepts

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# 1. Acids and Bases

## 1.1. Acids

An acid is a substance that donates hydrogen ions ($H^+$) when dissolved in water. For example, hydrochloric acid is a substance acid HCl in its aqueous solution (aq) dissociates as:

$$HCl\,(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Some examples of acids are:

(i) Hydrochloric acid (HCl) in gastric juice

(ii) Carbonic acid ($H_2CO_3$) in soft drinks

(iii) Ascorbic acid (vitamin C) in lemon and many fruits

## 1.2. Bases

A base is a substance that donates hydroxide ions ($OH^-$) when dissolved in water. For example, sodium hydroxide NaOH (aq) in its aqueous solutions dissociates as:

$$NaOH\ (aq) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

The term '**alkali**' is often used for water-soluble bases.

Some examples of bases are:

(i) Sodium hydroxide (NaOH) or caustic soda used in washing soaps

(ii) Potassium hydroxide (KOH) or potash used in bathing soaps

(iii) Ammonium hydroxide ($NH_4OH$) used in hair dyes

## 1.3 Indicators

There are many substances that show one colour in an acidic medium and another colour in a basic medium. Such substances are called **acid-base indicators**.

Litmus is a natural dye found in certain lichens. It was the earliest indicator to be used. It shows red colour in acidic solutions and blue colour in basic solutions. Phenolphthalein and methyl orange are some other indicators. The colours of these indicators in acidic, neutral and basic solutions are given below in table 1

## Table 1 Colours of some indicators in acidic and basic solutions

| Indicator | Colour in acidic solutions | Colour in neutral solutions | Colour in basic solutions |
|---|---|---|---|
| Litmus | red | purple | blue |
| Phenolphthalein | colourless | colourless | pink |
| Methyl orange | red | orange | yellow |

## 2. Properties of acids and bases

### Properties of Acids

The following are the characteristic properties of acids:

**1. Taste**

You must have noticed that some of the food items we eat have sour taste. The sour taste of many unripe fruits, lemon, vinegar and sour milk is caused by the acids present in them. Hence, we can say that acids have a sour taste. This is particularly true of dilute acids (see table 2).

**Table 2 Acids present in some common substances**

| Substance | Acid present |
|---|---|
| 1. Lemon juice | Citric acid and ascorbic acid (vitamin C) |
| 2. Vinegar | Ethanoic acid (commonly called acetic acid) |
| 3. Tamarind | Tartaric acid |
| 4. Sour milk | Lactic acid |

## 2. Conduction of electricity and dissociation of acids

Such solutions are commonly used in car and inverter batteries. When acids are dissolved in water, they produce ions which help in conducting the electricity. This process is known as dissociation. More specifically, acids produce hydrogen ions ($H^+$) which are responsible for all their characteristic properties. These ions do not exist as $H^+$ in the solution but combine with water molecules as shown below:

$$H^+ + H_2O \longrightarrow H_3O^+$$

**On the basis of the extent of dissociation occurring in their aqueous solutions, acids are classified as strong and weak acids:**

## Strong and Weak acids

Acids are classified as strong and weak acids and their characteristics are as follow:

| Strong Acids | Weak Acids |
|---|---|
| The acids which completely dissociate in water are called strong acids.<br>Nitric acid completely dissociates in water<br>$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$<br>There are only seven strong acids<br>1. HCl Hydrochloric Acid<br>2. HBr Hydrobromic Acid<br>3. HI Hydroiodic Acid<br>4. $HClO_4$ Perchloric Acid<br>5. $HClO_3$ Chloric Acid<br>6. $H_2SO_4$ Sulphuric Acid<br>7. $HNO_3$ Nitric Acid | The acids which dissociate partially in water are called weak acids. All organic acids like acetic acid and some inorganic acids are weak acids. Since their dissociation is only partial, it is depicted by double half arrows.<br>$HF(aq) \quad H^+(aq) + F^-(aq)$<br>The double arrows indicate here that<br>(i) the aqueous solution of hydrofluoric acid not only contains $H^+$ (aq) and $F^-$ (aq) ions but also the undissociated acid HF (aq).<br>(ii) there is an equilibrium between the undissociated acid HF (aq) and the ions furnished by it, $H^+$(aq) and $F^-$ (aq)<br>Examples:<br>(a) $CH_3COOH$ Ethanoic (acetic) acid<br>(b) HF Hydrofluoric acid<br>(c) HCN Hydrocynic acid<br>(d) $C_6H_5COOH$ Benzoic acid |

## 3. The reaction of acids with metals

For example, the reaction between zinc and dil. sulphuric acid can be written as:

$$\underset{\text{Zinc}}{Zn} + \underset{\text{dil. sulphuric acid}}{H_2SO_4} \rightarrow \underset{\text{zinc sulphate}}{ZnSO_4} + \underset{\text{hydrogen gas}}{H_2\uparrow}$$

## 4. The reaction of acids with metal carbonates and hydrogen carbonates

$$\underset{\text{sodium carbonate}}{Na_2CO_3(s)} + \underset{\text{dil. hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{sodium chloride}}{2NaCl(aq)} + \underset{\text{water}}{H_2O(l)} + \underset{\text{carbon dioxide}}{CO_2(g)\uparrow}$$

## 5. The reaction of acids with metal oxides

$$CuO(s) + 2HCl(aq) \longrightarrow CuCl_2(aq) + H_2O$$

## 6. Reaction of acids with bases

**Neutralization**. It results in the formation of salt and water. The reaction between hydrochloric acid and sodium hydroxide forms sodium chloride and water.

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O$$

## 8. Corrosive Nature

The ability of acids to attack various substances like metals, metal oxides and hydroxides is referred to as their corrosive nature. (It may be noted here that the term '**corrosion**' is used with reference to metals and refers to various deterioration processes (oxidation) they undergo due to their exposure to the environment). Acids are corrosive in nature as they can attack a variety of substances.

$$\underset{\text{(In glass)}}{\underset{\text{Silica}}{SiO_2}} + \underset{\text{acid}}{\underset{\text{hydrofluoric}}{4HF}} \longrightarrow \underset{\text{tetra fluoride}}{\underset{\text{silicon}}{SiF_4}} + \underset{\text{water}}{2H_2O}$$

# Properties of Bases

The following are the characteristic properties of bases:

**1. Taste and touch**

Bases have a bitter taste and their solutions are soapy to touch.

**2.Action on Indicators**

The colours shown by three commonly used indicators in presence of bases are listed below for easy recall.

**Table 3 Colours of some common indicators in basic solution**

| Indicator | Colour in basic medium |
|---|---|
| 1. Litmus | Blue |
| 2. Phenolphthalein | Pink |
| 3. Methyl orange | Yellow |

## 3. Conduction of electricity and dissociation of bases

Aqueous solutions (solution in water) of bases conduct electricity which is due to the formation of ions. Like acids, bases also dissociate on dissolving in water. Bases

produce hydroxyl ions ($OH^-$) which are responsible for their characteristic properties. The bases which are soluble in water and give $OH^-$ ions in their aqueous solution are called **alkalies**. All alkalies are bases but all bases are not alkalies. On the basis of the extent of dissociation occurring in their solution, bases are classified as strong and weak bases.

## Strong and Weak Bases

Bases are classified as strong and weak bases and their characteristics are as follow:

| **Strong Acids** | **Weak Acids** |
|---|---|
| These bases are completely dissociated in water to form the cation and hydroxide ion ($OH^-$). For example, potassium hydroxide dissociates as<br>$KOH\ (aq) \rightarrow K^+(aq) + OH^-(aq)$<br>There are only eight strong bases. These are the hydroxides of the elements of Groups 1 and 2 of the periodic table.<br>1. $LiOH$ Lithium hydroxide<br>2. $NaOH$ Sodium hydroxide<br>3. $KOH$ Potassium hydroxide<br>4. $RbOH$ Rubidium hydroxide<br>5. $CsOH$ Caesium hydroxide<br>6. $Ca\ (OH)_2$ Calcium hydroxide<br>7. $Sr\ (OH)_2$ Strontium hydroxide<br>8. $Ba\ (OH)_2$ Barium hydroxide | Weak bases do not furnish $OH^-$ ions by dissociation. They react with water to furnish $OH^-$ ions.<br>$NH_3(g) + H_2O$ $NH_4OH$<br>$NH_4OH\ (aq)$ $NH_4^+(aq) + OH^-(aq)$<br>or<br>$NH_3(g) + H_2O$ $NH^{4+}(aq) + OH^-(aq)$<br>The reaction resulting in the formation of OH– ions does not go to completion and the solution contains a relatively low concentration of OH– ions. The two half arrows are used in the equation to indicate that equilibrium is reached before the reaction is completed. Examples of weak bases (i) $NH_4OH$, (ii) $Cu(OH)_2$ (iv) $Cr(OH)_3$ (v) $Zn(OH)_2$ etc. |

## 4. Reaction of bases with metals

For example, sodium hydroxide reacts with zinc as shown below:

$$Zn(s) + 2NaOH(aq) \longrightarrow Na_2ZnO_2(aq) + H_2(g)\uparrow$$

## 5. Reaction of Bases with non-metal oxides

Bases react with oxides of non-metals like $CO_2$, $SO_2$, $SO_3$, $P_2O_5$ etc. to form salt and water. For example,

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O$$

## 6. Reaction of bases with acids

Such reactions are called neutralization reactions and result in the formation of salt and water. The following are some more examples of neutralization reactions:

$$HCl(aq) + KOH(aq) \longrightarrow KCl(aq) + H_2O$$

## Self-dissociation of water

Water plays an important role in acid-base chemistry. We have seen that it helps in the dissociation of acids and bases resulting in the formation of $H^+$(aq) and $OH^-$ (aq) ions respectively. Water itself undergoes a dissociation process which is called 'self-dissociation of water'.

$$H_2O \rightleftharpoons H^+(aq) + OH^-(aq)$$

The dissociation of water is extremely small and only about two out of every billion ($10^9$) water molecules are dissociated at 25°C. As a result, the concentrations of $H^+$(aq) and $OH^-$(aq) ions formed are also extremely low. At 25°C (298K):

$$[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ mol L}^{-1}$$

It must be noted here that in pure water and in all aqueous neutral solutions,

$$[H^+] = [OH^-]$$

Also, in pure water as well as in all aqueous solutions at a given temperature, product of concentrations of $H^+$(aq) and $OH^-$(aq) always remains constant. This product is called 'ionic product of water' and is given the symbol $K_w$. It is also called ionic product constant of water. Thus

$$K_w = [H^+]\,[OH^-]$$

At 25°C (298 K), in pure water, *Kw* can be calculated as:

$$K_w = (1.0 \times 10^{-7}) \times (1.0 \times 10^{-7})$$
$$= 1.0 \times 10^{-14}$$

**Table 3 Concentration of $H^+$(aq) ions in different types of aqueous solutions**

| Nature of solution | Concentration of $H^+$ ions at 25°C (298 K) |
|---|---|
| Neutral | $[H^+] = 1.0 \times 10^{-7}$ mol $L^{-1}$ |
| Acidic | $[H^+] > 1.0 \times 10^{-7}$ mol $L^{-1}$ |
| Basic | $[H^+] < 1.0 \times 10^{-7}$ mol $L^{-1}$ |

# pH and its importance

When dealing with a range of concentrations (such as these of $H^{+}$(aq) ions) that span many powers of ten, it is convenient to represent them on a more compressed logarithmic scale. We use the pH scale for denoting the concentration of hydrogen ions. pH notation was devised by the Danish biochemist Soren Sorensen in 1909. The term pH means "power of hydrogen". The pH is the logarithm of the reciprocal of the hydrogen ion concentration. It is written as:

$$pH = \log \frac{1}{[H^{+}]}$$

Alternately, the pH is the negative logarithm of the hydrogen ion concentration i.e

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

Because of the negative sign in the expression, if $[H^{+}]$ increases, pH would decrease and if it decreases, pH would increase.

**In pure water at 25° (298 K)**

$$[H^+] = 1.0 \times 10^{-7} \text{ mol L}^{-1}$$

$$\log[H^+] = \log(10^{-7}) = -7$$

$$\text{and pH} = -\log[H^+] = -(-7)$$

$$\text{pH} = 7$$

Since in pure water at 25°C (298 K)

$$[OH^-] = 1.0 \times 10^{-7} \text{ mol L}^{-1}$$

$$\text{Also, pOH} = 7$$

$$\text{Since, } K_w = 1.0 \times 10^{-14}$$

$$pK_w = 14$$

The relationship between $pK_w$, pH and pOH is

$$\mathbf{pK_w = pH + pOH}$$

at 25°C (298 K)

$$14 = \text{pH} + \text{pOH}$$

# pH Scale

The pH scale ranges from 0 to 14 on this scale. pH 7 is considered neutral, below 7 acidic and above 7 basic. The scale is shown below in Fig. 5.

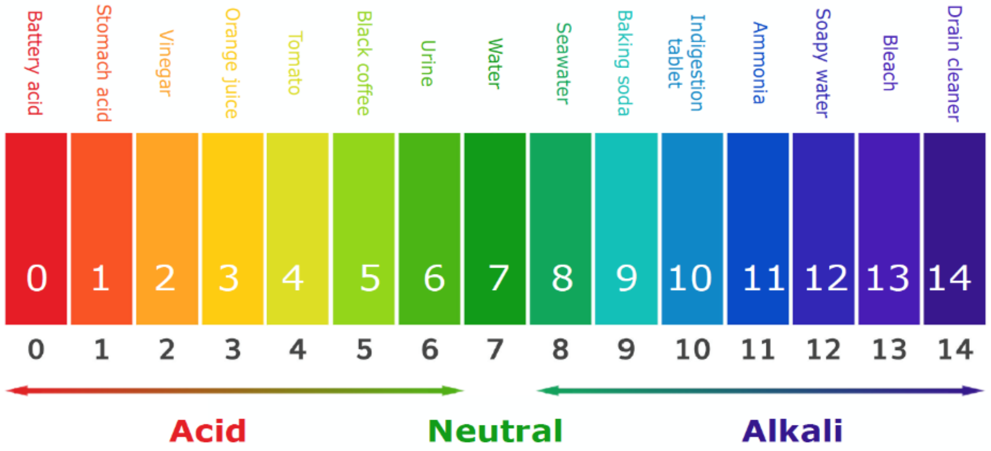


**Table 5: pH of some common acids and bases**

| Common Acids | pH | Common Bases | pH |
|---|---|---|---|
| HCl (4%) | 0 | Blood plasma | 7.4 |
| Stomach acid | 1 | Egg white | 8 |
| Lemon juice | 2 | Sea water | 8 |
| Vinegar | 3 | Baking soda | 9 |
| Oranges | 3.5 | Antacids | 10 |
| Soda, grapes | 4 | Ammonia water | 11 |
| Sour milk | 4.5 | Lime water | 12 |
| Fresh milk | 5 | Drain cleaner | 13 |
| Human saliva | 6-8 | Caustic soda 4% (NaOH) | 14 |
| Pure water | 7 | | |

# Calculations based on the pH concept

## A. Strong Acids and Bases

When the solution contains just one substance which is strong acid or base, the equation must be written in the ionic style and calculate the pH directly from the Hydrogen ion concentration (Hydronium ion).

**Example: Find the pH of a 0.03 M solution of hydrochloric acid, HCl**.

**Solution:** there is one substance in this solution which is HCl strong acid, this acid will completely dissociate in water to form $H^+$ and $Cl^-$.

$$HCl \rightarrow H^+(aq) + Cl^-(aq)$$

$$[H^+] = 0.03\ M$$

$$\mathbf{pH = -log[H^+]} = 1.5$$

**Example: Find the pH of a 0.03 M solution of NaOH.**

**Solution:** there is one substance in this solution which is NaOH strong Base, this the base will completely dissociation in water to form $OH^-$ and $Na^+$

$$NaOH \rightarrow Na^+(aq) + OH^-(aq)$$

$$[\mathbf{OH^-}] = 0.03\ M$$

$$pOH = -\log[\mathbf{OH^-}] = \mathbf{1.5}$$

$$pH = 14 - pOH = 12.5$$

## A. Weak Acids and Bases

### Weak Acids and Acid Ionization Constants

Weak acids do not fully dissociate in water. The extent to which a weak acid ionizes in water depends on:

1. The concentration of the acid

2. The equilibrium constant for the ionization reaction, $K_a$

We will only consider monoprotic acids in this section:

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

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

$K_a$ is called the acid ionization constant.

$$pK_a = -\log K_a$$

We can calculate pH from Ka, or Ka from pH.

**Example: Calculate the pH of a 0.50 M HF solution at 25°C? $K_a$ = 7.1 x $10^{-4}$**

| | $HF \rightleftharpoons$ | $H^+$ + | $F^-$ |
|---|---|---|---|
| Initial | 0.50 | 0* | 0 |
| Change | -x | +x | +x |
| Equilibrium | 0.50 - x | x | x |

we have neglected the auto-ionization of water because it will have a negligible effect (x will be much larger than $10^{-7}$)

At equilibrium we have

$K_a$ = $x^2$ /0.50-x

This is a quadratic equation in x. The “simple” math is to put

(0.50 − x) = 0.50

Then we have

$K_a$ = $x^2$/0.50

$x^2$ = (0.50)×(7.1 x $10^{-4}$)

x = 0.019

The pH is then = -log(0.019) = 1.72

# Using pH to Determine $K_a$

We can use the pH of a weak acid to determine the equilibrium concentrations, which gives us the equilibrium constant.

**Example**: Calculate the $K_a$ of a weak acid if a 0.015 M solution of the acid has a pH of 5.03 at 25ºC.

**Solution:**

pH = 5.03 = −log[H⁺]

$[H^+] = 10^{-5.03} = 9.33 \times 10^{-6}$ M

| | $HA \rightleftharpoons$ | $H^+$ + | $A^-$ |
|---|---|---|---|
| Initial | 0.015 | 0 | 0 |
| Change | $- 9.33 \times 10^{-6}$ | $+ 9.33 \times 10^{-6}$ | $+ 9.33 \times 10^{-6}$ |
| Equilibrium | 0.01499 | $9.33 \times 10^{-6}$ | $9.33 \times 10^{-6}$ |

$K_a = (9.33 \times 10^{-6})^2/0.01499 = 5.8 \times 10^{-9}$

# Calculating pH from $K_b$

We proceed as for weak acid, using pH + pOH = 14

**Example**: Calculate the pH at 25ºC of a 0.16 M solution of a weak base with a $K_b$ of 2.9 x $10^{-11}$.

**Solution**: Since $K_b$ is so small we can make the math simple.

| | B (+ $H_2O$) | $\rightleftharpoons$ $HB^+$ + | $OH^-$ |
|---|---|---|---|
| Initial | 0.16 | 0 | 0 |
| Change | -x | +x | +x |
| Equilibrium | 0.16 - x | x | x |

$K_b$ =[$HB^+$][$OH^-$]/[B]

= $x^2/0.16 - x$

= $x^2/0.16$

Which gives $x^2 = (0.16) \times K_b$

x = 2.15 x $10^{-6}$

pOH = 5.67

Therefore pH = 14 – 5.67 = 8.33

# Using pH to Determine $K_b$

This is very similar the weak acid calculation.

**Example:** Determine the $K_b$ of a weak base if a 0.35 M solution of the base has a pH of 11.84 at 25ºC.

**Solution**:

pOH = 14 – pH = 2.16 ⇒ $[OH^-] = 10^{-2.16} = 6.92 \times 10^{-3}$

| | B (+ $H_2O$) | $\rightleftharpoons$ $HB^+$ + | $OH^-$ |
|---|---|---|---|
| Initial | 0.35 | 0 | 0 |
| Change | $- 6.92 \times 10^{-3}$ | $+ 6.92 \times 10^{-3}$ | $+ 6.92 \times 10^{-3}$ |
| Equilibrium | 0.3431 | $6.92 \times 10^{-3}$ | $6.92 \times 10^{-3}$ |

$K_b = (6.92 \times 10^{-3})^2 / 0.3431$

$= 1.4 \times 10^{-4}$

# Salts

A salt is an ionic compound consisting of a cation except a hydrogen ion and an anion except a hydroxide ion

– A salt is the product of acid-base neutralization

– In general acid + base → salt + water

$$NaOH + H_2CO_3 \leftrightarrow NaHCO_3 + H_2O$$

**Characteristics of salt:**

- Most of the salts are crystalline solid.
- Salts may be transparent or opaque.
- Most of the salts are soluble in water.
- Solution of the salts conducts electricity in their molten state also.
- The salt may be salty, sour, sweet, bitter and umami (savoury).
- Neutral salts are odourless.
- Salts can be colourless or coloured.

# Types of salts

**(i) Neutral Salt:** Salts produced because of reaction between a strong acid and strong base are neutral in nature. The pH value of such salts is equal to 7, neutral. Example: Sodium chloride, Sodium sulphate. Postassium chloride, etc.

**Sodium chloride (NaCl):** It is formed after the reaction between hydrochloric acid (a strong acid) and sodium hydroxide (a strong base)

$$NaOH + HCl \rightarrow NaCl + H_2O$$

**(ii) Acidic Salts:** Salts which are formed after the reaction between a strong acid and weak base are called Acidic salts. The pH value of acidic salt is lower than 7. For example, Ammonium sulphate, Ammonium chloride, etc.

Ammonium chloride is formed after reaction between hydrochloric acid (a strong acid) and ammonium hydroxide (a weak base).

$$NH_4OH + HCl \rightarrow NH_4Cl + H_2O$$

**(iii) Basic Salts:** Salts that are formed after the reaction between a weak acid and strong base are called Basic Salts. For example; Sodium carbonate, Sodium acetate, etc.

Sodium carbonate is formed after the reaction between sodium hydroxide (a strong base) and carbonic acid (a weak acid).

$$H_2CO_3 + 2NaOH \rightarrow Na_2CO_3 + H_2O$$