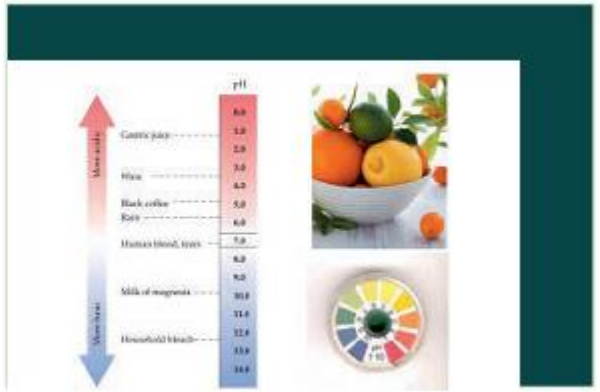
UNIT 1



ACID-BASE EQUILIBRIA

Prepared by:Z Secret Training Institute

Main Content of the Unit

1.1 Acid-Base Concepts

1.2 Ionic Equlibria of Week Acids and Bases

1.3 Common Ion Effect and Buffer Solution

1.4 Hydrolysis of Salts

1.5 Acid–Base Indicators and Titrations

Introduction

- The concepts of acids and bases are probably among the most familiar chemistry concepts.
- The reason is that acids and bases have been used as laboratory chemicals for centuries, as well as in the home.
- Common household acids include acetic acid (vinegar), citric acid, in citrus fruits), and phosphoric acid (a flavoring in carbonated beverages).
- Sodium hydroxide (NaOH , drain cleaner) and ammonia (NH3 , glass cleaner), are household bases.

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1.1 ACID-BASE CONCEPTS

At the end of this subunit, you will be able to:

- Identify the acid-base conjugate pairs from the given reaction;
- Write an equation for self-ionization of water and ammonia.
- Explain what is meant by amphiprotic species
- Define acid and base by the Lewis concept
- Define acid and base by the Bronsted-Lowry concept
- Define acid and base by the Arrhenius Concept
- Explain what conjugate acids and conjugate bases are

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1. Arrhenius Concept of Acids and Bases

- The Swedish chemist Svante Arrhenius framed the first successful concept of acids and bases.
- He defined acids and bases in terms of the effect these substances have on water.

❑ **According to Arrhenius,**

- Acids are substances that increase the concentration of H^+ (proton ion) in aqueous solution
- Bases increase the concentration of OH^- (a hydroxide ion) in aqueous solution.

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- In Arrhenius's theory,

- A strong acid is a substance that completely ionizes in aqueous solution to give H_3O^+ (aq) and an anion.
- An example is perchloric acid, $HClO_4$

$$HClO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ClO_4^-(aq)$$

- A strong base completely ionizes in aqueous solution to give *OH* and a cation.
- Sodium hydroxide is an example of a strong base.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

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The Arrhenius definition of acids and bases has the following limitations.

- Chemistry occurs in many reactions that take place in solvents other than water, and this definition does not work for such reactions.
- It does not explain why some compounds that contain hydrogen such as HCl dissolve in water to form acidic solution where as NH_3 and CH_4
- Only the compounds that contain the OH^- ion can be classified as Arrhenius bases. The Arrhenius definition does not explain why compounds such as Na_2CO_3 have basic properties.

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Exercise 1.1

1. Based on their dissociations in water solution, classify each of the following compounds as Arrhenius acid, Arrhenius base, or as a compound that cannot be classified as an Arrhenius acid or Arrhenius base.

a. $HBr(aq) + H_2O(l) \rightarrow H_3O^+(aq) + 2Br^-(aq)$

b. $NaCl(s) + H_2O(l) \rightarrow Na^+(aq) + Cl^-(aq)$

c. $NaOH(aq) + H_2O(l) \rightarrow Na^+(aq) + 2OH^-(aq)$

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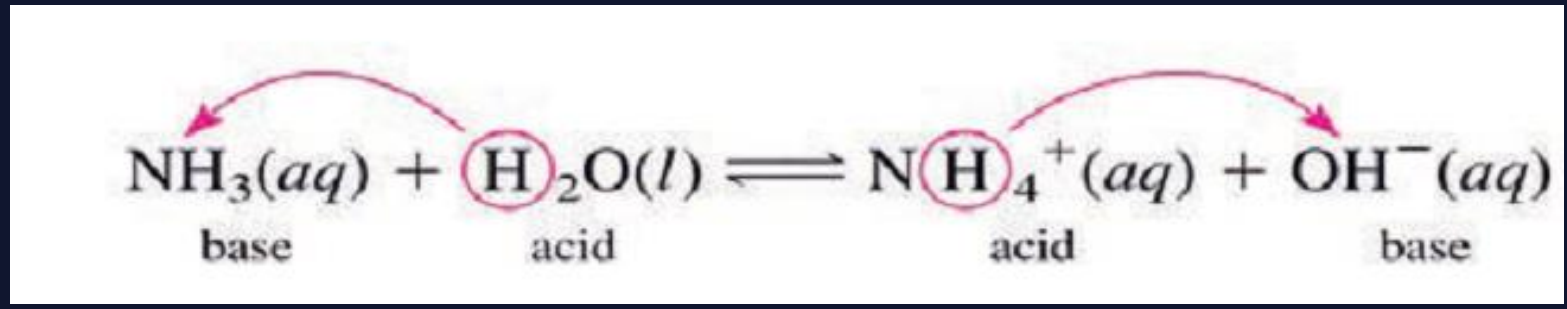
2 . Brønsted-Lowry Concept of Acids and Bases

- In 1923, J. N. Brønsted in Denmark and T. M. Lowry in Great Britain independently proposed a new acid base theory.
- They pointed out that acid–base reactions can be seen as proton- transfer reactions and those acids and bases can be defined in terms of proton (H) transfer.
- According to their concept, an acid is a proton donor and a base is a proton acceptor.

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Consider the ionization of hydrochloric acid in water:





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❑ Notes:

- All Arrhenius acids are bronsted-lowery acids but not the reverse
- All Arrhenius bases contain hydroxide ion which is bronsted –lowery base.
- But all bronsted – lowery base are not Arrhenius bases

Weakness of bronsted-lowery theory

- It is restricted an acid base concept only to proton transfer
- It doesn’t explain why species like CO_2 ,SO_2 ,Cu^{2+} function as acid even if they don’t contain proton

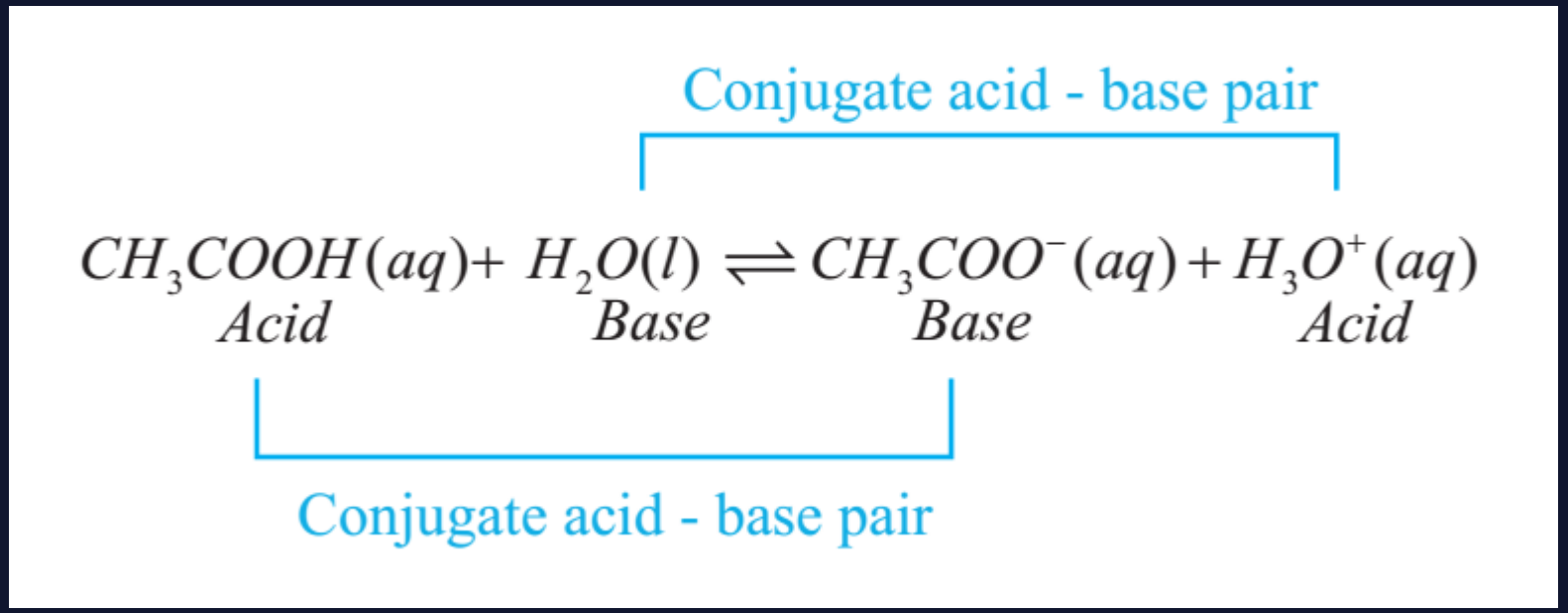
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- The advantage of this definition is that it is not limited to aqueous solutions.
- Bronsted-Lowry acids and bases always occur in pairs called conjugate acid base pairs.
- Conjugate acid-base pairs can be defined as an acid and its conjugate base or a base and its conjugate acid.

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- **The conjugate base of a Brønsted-Lowry acid** is the species that remains when one proton has been removed from the acid.
- **Conjugate acid** results from the addition of a proton to a Brønsted-Lowry base.

❑ Every Brønsted-Lowry acid has a conjugate base, and every Brønsted- Lowry base has a conjugate acid.



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For any conjugate acid-base pair

- The conjugate base has one *fewer H* and one *more* minus charge than the acid.
- The conjugate acid has one *more H* and one more + charge than the base.



HF	+	H_2O	F^-	+	H_3O^+
HCOOH	+	CN^-	$HCOO^-$	+	HCN
NH_4^+	+	CO_3^{2-}	NH_3	+	HCO_3^-

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Exercise 1.2

1. Identify the Brønsted-Lowry acids, bases, conjugate acids and conjugate bases in each of the following reaction

   a. $HClO_2 + H_2O \rightleftharpoons ClO_2^- + H_3O^+$

   b. $OCl^- + H_2O \rightleftharpoons HOCl + OH^-$

   c. $H_2O + SO_3^{2-} \rightleftharpoons OH^- + HSO_3^-$

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Strengths of Conjugate Acid-Base Pairs

- The stronger the acid, the weaker is its conjugate base.
- Similarly, the stronger the base, the weaker is its conjugate acid
- **For example:** *HCl* is a strong acid, and its conjugate base Cl^-, is a weak base.

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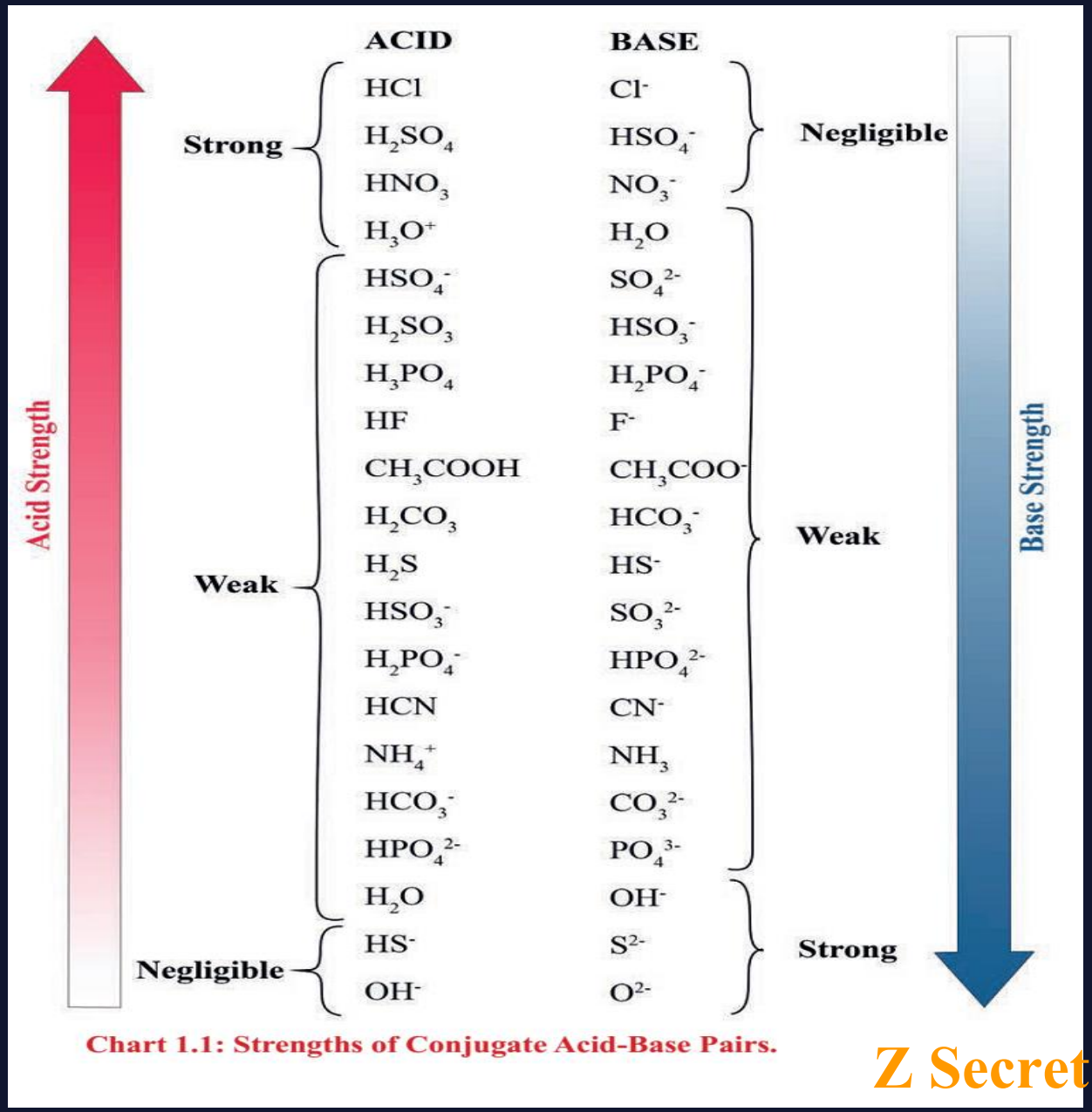


Chart 1.1: Strengths of Conjugate Acid-Base Pairs.

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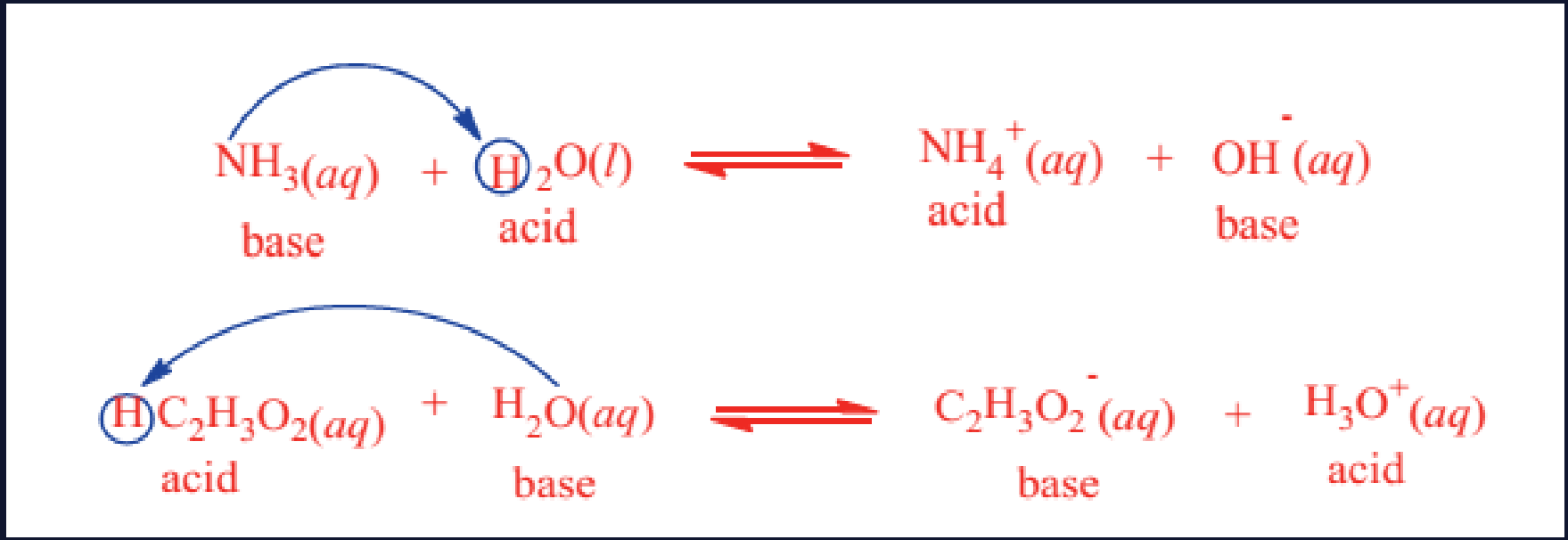
Auto Ionization of Substances

- **Molecular auto ionization (or self-ionization)** is a reaction between two identical neutral molecules, especially in a solution, to produce an anion and a cation.
- If a pure liquid partially dissociates into ions, it is said to be self-ionizing.

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Amphiprotic Species

- Molecules or ions that can either donate or accept a proton, depending on the other reactant, are called **amphiprotic species**

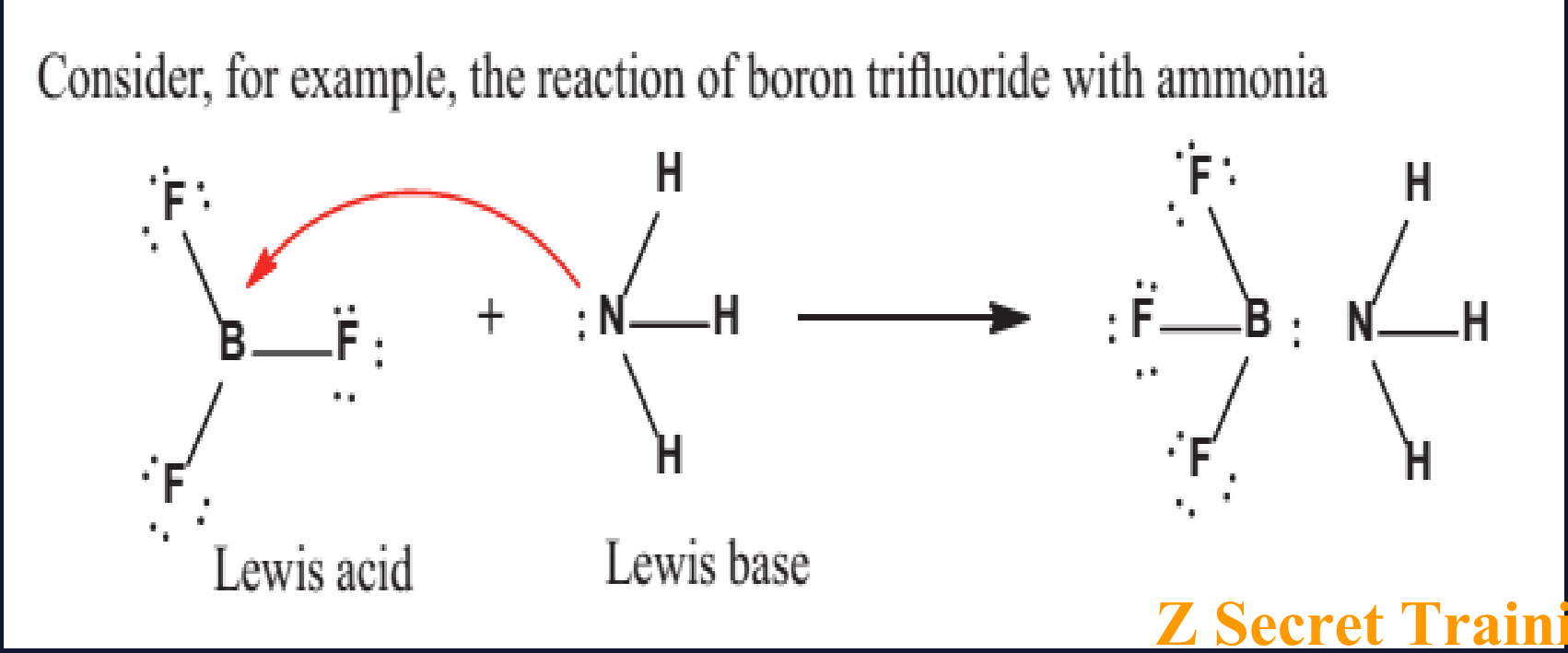


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3. Lewis Concept of Acids and Bases

❑ **According to this concept, the Lewis acid-base definition holds that**

- **A base is any species that donates an electron pair to form a bond.**
- **An acid is any species that accepts an electron pair to form a bond.**



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❑ Note:

- Any molecule or negatively charged species having an excess of electrons can be considered as a **Lewis base**
- Any electron-deficient molecule or positively charged species can be considered as a **Lewis acid**.

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Exercise 1.4

1. Identify Lewis acids and Lewis bases in each of the following reactions.

   a. $H^+ + OH^- \rightleftharpoons H_2O$

   b. $Cl^- + BCl_3 \rightleftharpoons BCl_4^-$

   c. $K^+ + 6H_2O \rightleftharpoons K(H_2O)_6^+$

   d. $OH^- + Al(OH)_3 \rightleftharpoons Al(OH)_4^-$

   e. $CO_2 + H_2O \rightleftharpoons H_2CO_3$

   f. $Ni + 4CO \rightleftharpoons Ni(CO)_4$

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1.2 IONIC EQUILIBRIA OF WEAK ACIDS AND BASES

❑ **At the end of this subunit, you will be able to:**

- describe the ionization of water
- derive the expression of ion product for water, Kw
- explain why water is a weak electrolyte
- use Kw to calculate [H3O+] or [OH-] in aqueous solution
- write an expression for the percent ionization of weak acids or weak bases
- calculate the percent dissociation of weak acids and bases
- write the expression for the acid-dissociation constant, Ka
- calculate Ka for an acid from the concentration of a given solution and it's pH

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- calculate $[H^+]$ and pH of an acidic solution from given values of Ka and the initial concentration of the solution
- write the expression for the base-dissociation constant, Kb ;
- calculate Kb for a base from the concentration of a basic solution and its pOH
- calculate the $[OH^-]$ and pOH of a basic solution from a given value of Kb and the initial concentration of the solution.

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1.2 IONIC EQUILIBRIA OF WEAK ACIDS AND BASES

Ionization of Water

- **How do you calculate the concentration of H_3O^+ ions if the concentrations of OH ions and Kw at 25°C are given?**
- Although pure water is often considered a non-electrolyte (nonconductor of electricity), precise measurements do show a very small conduction.
- This conduction results from self-ionization (or auto ionization) of water, a reaction in which two like molecules react to give ions.
- The H_2O molecule can act as either an acid or a base; it is amphiprotic.

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- Like any equilibrium process, the auto-ionization of water is described quantitatively by an equilibrium constant:

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

- Because the concentration of ions formed is very small and the concentration of H_2O remains essentially constant, about 56M at 25°C, we multiply Kc by $[H_2O]^2$ to obtain a new equilibrium constant, the ion-product constant for water, Kw :

$$K_c[H_2O]^2 = [H_3O^+][OH^-]$$

- We call the equilibrium value of the ion product $[H_3O^+][OH^-]$, the ion-product constant for water, Kw

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- At 25 ^{o}C, the value of Kw is 1.0 x 10^{-14}.
- Like any equilibrium constant, Kw varies with temperature.
- At body temperature (37^{o}C), Kw equals 2.5 x 10^{-14} .
- Using Kw , you can calculate the concentrations of [H_3O^+] and [OH^-]ions in pure water

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- If you add an acid or a base to water, the concentrations of $[H_3O^+]$ and $[OH^-]$ will no longer be equal. The equilibrium-constant equation $K_w = [H_3O^+][OH^-]$ will still hold.
- In any aqueous solution at 25°C, no matter what it contains, the product of $[H_3O^+]$ and $[OH^-]$ must always equal 1.0×10^{-14}
- The equilibrium nature of auto-ionization allows us to define "acidic" and "basic" solutions in terms of relative magnitudes of $[H_3O^+]$ and $[OH^-]$.

A neutral solution, where $[H_3O^+] = [OH^-]$.

An acidic solution, where $[H_3O^+] > [OH^-]$.

A basic solution, where $[OH^-] > [H_3O^+]$.

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Example 1.3

A chemistry researcher adds a measured amount of HCl gas to pure water at 25 ^{0}C and obtains a solution with $\left[H_3O^+\right]$= 3.0 x 10^{-4} M. Calculate $\left[OH^-\right]$. Is the solution neutral, acidic, or basic?

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Exercise 1.5

1. Calculate $[H^+]$ or $[OH^-]$, as required, for each of the following solutions at 25°C,and state whether the solution is neutral, acidic, or basic.

   a. $[H^+] = 1.0 \times 10^{-7} M$ b. $[OH^-] = 1.0 \times 10^{-4} M$

   c . $[OH^-] = 1.0 \times 10^{-8} M$

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2. Calculate the concentration of OH^- in a solution in which

a. $\left[H_3O^+\right] = 2.0 \text{ x } 10^{-5} M$ b. $\left[H_3O^+\right] = \left[OH^-\right]$

c. $\left[H_3O^+\right] = 10^2 \ x \ \left[OH^-\right]$

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3. Calculate $\left[H_3O^+\right]$ in a solution that is at 25 °C and has $\left[OH^-\right] = 6.7 \times 10^{-2} M$. Is the solution neutral, acidic, or basic?

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4. Why water is a weak electrolyte?

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The pH scale

- It is a well-known fact that whether an aqueous solution is acidic, neutral, or basic depends on the hydronium-ion concentration.
- You can quantitatively describe the acidity by giving the hydronium-ion concentration.
- But because these concentration values may be very small, it is often more convenient to give the acidity in terms of pH

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- pH is a measure of the hydronium ion content of a solution
- The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration (in mol/L):

$$pH = -\log [H_3O^+] \text{ or } pH = -\log [H^+]$$

- Note: that the negative logarithm gives us positive numbers for pH.
- Like the equilibrium constant, the pH of a solution is a dimensionless quantity.

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The *pH* of a solution is measured by a *pH* - meter.



Figure 1.1:pH meter.

Activate W
Go to Setting

- pH meters have scales marked with values from 1 to 14, *pH* values can, in fact, be greater than 1 and less than 14.
- pH value decreases as the concentration of H^+ ions increases; in other words, the more acidic the solution, the lower its *pH* ;
- The more basic the solution, the higher its *pH* .
- For a neutral solution, $pH = 7$
- Acidic solutions have $pH < 7$
- Basic solutions have $pH > 7$

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pOH

- A pOH scale analogous to the pH scale can be devised using the negative logarithm of the hydroxide ion concentration of a solution. Thus, we define pOH as:

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

- If we are given the pOH value of a solution and asked to calculate the OH ion concentration, we can take the antilog of the above equation.

Relation between pH & pOH

- pH + pOH=14 @25℃
- Thus, in general, the sum of the pH and pOH values must equal pKw

Example 1.4

1. Calculate:

   a. the pH and pOH of a juice solution in which $\left[H_3O^+\right]$ is 5.0×10^{-3} M

   b. the $\left[H_3O^+\right]$ and $\left[OH^-\right]$ of human blood at $pH = 7.40$

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Exercise 1.6

1. A solution formed by dissolving an antacid tablet has a *pH* of 9.18 at 25 °C. Calculate $[H^+]$, $[OH^-]$ and pOH.

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2. A solution is prepared by diluting concentrated HNO_3 to 2.0 M, 0.30 M and 0.0063 M HNO_3 at 25 °C. Calculate $[H^+]$, $[OH^-]$, pH and pOH f the three solutions.

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Measures of the Strengths of Acids and Bases in Aqueous Solution

- **At the end of this subunit, you will be able to:**

- calculate the percent dissociation of weak acids and bases
- write the expression for the acid-dissociation constant, Ka
- calculate Ka for an acid from the concentration of a given solution and it's pH
- calculate $[H^+]$ and pH of an acidic solution from given values of Ka and the initial concentration of the solution
- write the expression for the base-dissociation constant, Kb ;
- calculate Kb for a base from the concentration of a basic solution and its pOH
- calculate the $[OH^-]$ and pOH of a basic solution from a given value of Kb and the initial concentration of the solution

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Measures of the Strengths of Acids and Bases in Aqueous Solution

- The concentration of hydrogen and hydroxide ions
- pH and pOH
- Percent Ionization
- percent dissociation, Ka and Kb.

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1. The concentration of hydrogen and hydroxide ions

- By definition a strong acid is one that completely dissociates in water to release proton.
- In solutions of the same concentration, stronger acids ionize to a greater extent, and so yield higher concentrations of hydronium ions (H_3O^+) than do weaker acids
- By contrast, however, a weak acid, being less willing to donate its proton, will only partially dissociate in solution

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- **Base strength refers to the ability of a base to accept protons.**
  - A strong base accepts more protons readily than a weak base.
  - A solution of a stronger base will contain a larger concentration of hydroxide ions than a solution of a weaker base if both solutions are of equal concentration

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2. pH and pOH

- One way to determine the strength of acids is using the *pH* values.
- The acid strength increase with smaller *pH* value,
- The concentration of hydroxide ions in a solution can be expressed in terms of the *pOH* of the solution.
- Hence, the strength of bases can also be determined from their pOH values.
- The strength of base increases with decreasing the pOH value.

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3. Percent Ionization

- Another measure of the strength of an acid and base is its percent ionization, which is defined as the proportion of ionized molecules on a percentage basis.

❑ **Mathematically,**

$$\text{Percent Ionization of Acid} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\ \%$$

$$\text{Percent Ionization of Base} = \frac{\text{Ionized base concentration at equilibrium}}{\text{Initial concentration of base}} \times 100\ \%$$

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3. Dissociation (Ionization) Constants

Acid ionization Constant, Ka

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

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Table 1.2: Ionization constant of some weak monoprotic acids at 25°C.

Name of the Acid	Formula	K_a
Acetic acid	CH_3COOH	1.8×10^{-5}
Ascorbic acid	$C_6H_8O_6$	8.0×10^{-5}
Benzoic Acid	C_6H_5COOH	6.5×10^{-5}
Formic acid	$HCOOH$	1.7×10^{-4}
Hydrocyanic acid	HCN	4.9×10^{-10}
Hydrofluoric acid	HF	6.8×10^{-4}
Hypobromous acid	$HOBr$	2.5×10^{-9}
Hypochlorous acid	$HOCl$	3.0×10^{-8}
Nitrous acid	HNO_2	4.5×10^{-4}

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Example 1.5

1. A 0.250 M aqueous solution of butyric acid is found to have a pH of 2.72. Determine K_a for butyric acid

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2. Calculate the pH of a 0.50 M HF solution at 25°C. The ionization of HF is given by $HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$

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3. What is the percent ionization of acetic acid in a 0.100M solution of acetic acid, CH_3COOH ?

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Exercise 1.7

1. Calculate the percent ionization of a 0.10M solution of acetic acid with a pH of 2.89.
2. Calculate the pH of a 0.10 M solution of acetic acid. $K_a = 1.8 \times 10^{-5}$
3. For a 0.036 M HNO_2 solution.
   a. Write a chemical equation that shows the ionization of nitrous acid in water.
   b. Calculate the equilibrium concentration of hydrogen ions and nitrous acid at 25°C, using the approximation method. Then check whether the approximation is valid or not.
   c. If the approximation is invalid, use the quadratic formula to calculate the concentration of hydrogenions.
   d. Calculate the pH of the solution.

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Base dissociation constant, Kb

- In general, a weak base B with the base ionization

$$B(aq) + H_2O(l) \rightleftharpoons HB^+(aq) + OH^-(aq)$$

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

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Table 1.3: shows the K_b values of some common weak bases at 25°C

Base	Formula	K_b
Ammonia	NH_3	1.8×10^{-5}
Aniline	$C_6H_5NH_2$	4.0×10^{-10}
Ethylamine	$C_2H_5NH_2$	4.7×10^{-4}
Hydrazine	C_2H_4	1.7×10^{-6}
Hydroxylamine	NH_2OH	1.1×10^{-6}
Methylamine	CH_3NH_2	4.4×10^{-4}
Pyridine	C_5H_5N	1.7×10^{-9}

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Example 1.6

1. What is the hydronium-ion concentration of a 0.20 *M* solution of ammonia in water? $K_b = 1.8 \times 10^{-5}$

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Exercise 1.8

1. For a 0.040 M ammonia **solution:**
   a. Write a chemical equation that shows the ionization of ammonia in water.
   b. Calculate the equilibrium concentration of ammonia, ammonium ions and hydroxide ions, using the approximation method. Check whether the approximation is valid or not.
   c. If the approximation is invalid, use the quadratic formula to calculate the concentration of ammonia.
   d. Calculate the pOH and pH of the solution.

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The Common Ion Effect and Buffer solution

Common-ion effect

- The common-ion effect is the shift in an ionic equilibrium caused by the addition of a solute that provides an ion that takes part in the equilibrium.
- The common-ion effect occurs when a given ion is added to an equilibrium mixture that already contains that ion.
- Consider a solution of acetic acid, CH3 COOH , in which you have the following acid ionization equilibrium:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

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- Suppose you add HCl (aq) to this solution. What is the effect on the acid-ionization equilibrium?

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

- Because HCl (aq) is a strong acid, it provides H_3O^+ ion, which is present on the right side of the equation for acetic acid ionization.
- According to LeChâtelier's principle, the equilibrium composition should shift to the left.
- The degree of ionization of acetic acid is decreased by the addition of a strong acid.

$$CH_3COOH(aq) + H_2O(l) \leftarrow CH_3COO^-(aq) + \underset{\text{added}}{H_3O^+(aq)}$$

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- Another example is, if sodium acetate and acetic acid are dissolved
- in the same solution, they both dissociate and ionize to produce CH_3COO^-ions, we can represent the effect of acetate salts on the acetic acid equilibrium as:

When salt supplies common anion CH_3COO^-, the equilibrium shifts to form more CH_3COOH

$$CH_3COONa(aq) \longrightarrow Na^+ + CH_3COO^-$$

Added CH_3COO^-

$$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^- \quad K_a = 1.8 \times 10^{-5}$$

Equilibrium shifts to form more CH_3COOH



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Example 1.7

1. Determine the $\left[H_3O^+\right]$ and $\left[CH_3COO^-\right]$ in a solution that is 0.10 M in both CH_3COOH and HCl.

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Exercise 1.9

1. Calculate the pH of a solution containing 0.20 M CH_3COOH and 0.30 M CH_3COONa
2. What would be the pH of a 0.20 M CH_3COOH solution if no salt were present?

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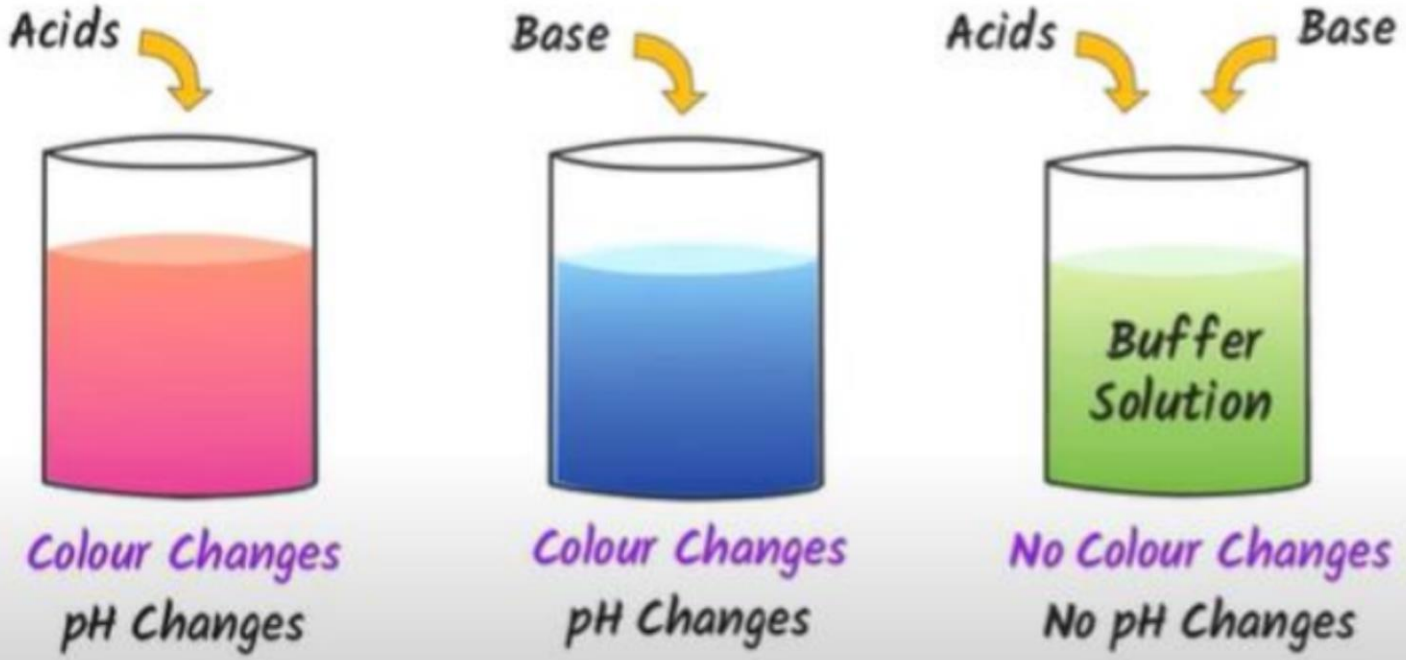
Buffer Solutions

- A buffer is commonly defined as a solution that resists changes in pH when a small amount of acid or base is added or when the solution is diluted with pure solvent.
- A buffer solution actually is a mixture of a weak acid and its conjugate base or a mixture of a weak base and its conjugate acid.
- The conjugate forms are commonly referred to as “salts”.

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Buffer Solutions

Basic Concept



Buffer Solution

"A solution which can resist change in pH when small amount of acids or bases is added."

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Table1.4: Comparison of buffered and un buffered solutions.

	Initial pH of 1.0 L sample	pH after addition of 0.010 mol NaOH	pH after addition of 0.010 mol HCl
Un buffered **solution:** 1.28×10^{-5} M *HCl*	4.8	12.0	2.0
Buffered **solution:** 0.099 M CH_3OOH 0.097 M CH_3COONa	4.8	4.8	4.7

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- A buffer solution must contain a relatively large concentration of acid to react with any OH^- ions that may be added to it.
- Similarly, it must contain a large concentration of base to react with any added H^+ ions.
- Furthermore, the acid and the base components of the buffer must not consume each other in a neutralization reaction.
- a buffer solution resists changes in pH through its ability to combine with the H^+ and OH^- ions.

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- How buffer solution resist change in pH

❑ Importance of buffer solution

- Buffers are very important to chemical and biological systems.
- The pH in the human body varies greatly from one fluid to another; for example, the pH of blood is about 7.4, whereas the gastric juice in our stomachs has a pH of about 1.5.
- These pH values, which are crucial for the proper functioning of enzymes and the balance of osmotic pressure, are maintained by buffers in most cases.

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- The pH of a buffer solution can be estimated with the help of **Henderson–Hasselbalch equation** when the concentration of the acid and its conjugate base, or the base and the corresponding conjugate acid, are known.

Henderson-Hasselbalch equation:

$$pH = pK_a + \log\left(\frac{[Conjugate\ base]}{[weak\ acid]}\right)$$

Similarly, for a weak base dissociation:

$$pOH = pK_b + \log\left(\frac{[Conjugate\ acid]}{[weak\ base]}\right)$$

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Example 1.8

1. What is the pH of a buffer solution consisting of 0.035M NH_3 and 0.050M NH_4^+ (K_a for NH_4^+ is 5.6×10^{-10})? The equation for the reaction is:

$$NH_4^+ \rightleftharpoons H^+ + NH_3$$

Assuming that the change in concentrations is negligible in order for the system to reach equilibrium, the Henderson-Hasselbalch equation will be:

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2. A buffer is made by mixing 0.060 M NH_3 with 0.040 M NH_4Cl. What is the pH of this buffer? $K_b = 1.8 \times 10^{-5}$

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Exercise 1.10

1. Calculate the pH :
   a. of a buffer solution containing 0.1 M CH_3COOH and a 0.1 M solution of CH_3COONa .
   b. when 1.0 mL of 0.10 M HCl is added to 100 mL of the buffer in (a);
   c. when 1.0 mL of 0.10 M $NaOH$ is added to 100 mL of the buffer in (a);
   d. of an unbuffered solution containing 1.8×10^{-5} HCl ;
   e. change of an unbuffered solution in (d) after adding
      i. 1.0 mL of 0.1 M $NaOH$ to 100 mL of the solution,
      ii. 1.0 mL of 0.10 M HCl to 100 mL of the solution.

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1.4 Hydrolysis of salt

At the end of this subunit, you will be able to:

- define hydrolysis
- explain why a salt of weak acid and strong base gives a basic solution
- explain why a salt of strong acid and weak base gives an acidic solution
- explain why salts of weak acids and weak bases give acidic, basic or neutral solutions

Hydrolysis of salt

- **Hydrolysis** is derived from a Greek word **hydro meaning water** and **lysis meaning break or to unbind**.
- **Hydrolysis** is a common form of a chemical reaction where water is mostly used to break down the chemical bonds that exist between particular substances.
- Usually in hydrolysis the water molecules get attached to two parts of a molecule.
- One molecule of a substance will get H^+ ion and the other molecule receives the OH^- group.
- **Example:**

$$CH_3COOCH_2CH_3 + H_2O \rightarrow CH_3COOH + CH_3CH_2OH$$

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- The term **salt hydrolysis** describes the interaction of anion and cation of a salt, or both, with water.
- Depending on the strengths of the parent acids and bases, the cation and anion of a salt can serves as an acid, base or neutral.

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Hydrolysis of Salts of Strong Acids and Strong Bases

- The reaction between a strong acid (say, HCl) and a strong base (say, NaOH) can be represented by

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H\ O(l)$$

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- The anions derived from strong acids are weak conjugate bases and do not undergo hydrolysis.
- The cations derived from strong bases are weak conjugate acids and also do not hydrolysis.
- For the above reaction, chloride ions, Cl^-, and sodium ions, Na^+ , do not hydrolyze, It involves only ionization of water and no hydrolysis.
- The solution of the salt will be neutral (pH=7).

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Hydrolysis of Salts of Weak Acids and Strong Bases

- Consider the neutralization reaction between acetic acid (a weak acid) and sodium hydroxide (a strong base):

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(aq)$$

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Hydrolysis of Salts of Strong Acids and Weak Bases

- For example, ammonium chloride, NH_4Cl , is a salt formed by the reaction of the weak base ammonia with the strong acid HCl :

$$NH_3(aq)+HCl(aq) \rightarrow NH_4Cl(aq)$$

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Hydrolysis of Salts of Weak Acids and Weak Bases

- Solutions of a salt formed by the reaction of a weak acid and a weak base involve both cationic and anionic hydrolysis.
- To predict whether the solution is acidic or basic you need to compare the Ka of the weak acid and the Kb of the weak base.
- If the Ka is greater than the Kb, the solution is acidic, and if the Ka is less than the Kb , the solution is basic.
- If they are equal the solution is neutral.
- Example consider solutions of Ammonium acetate

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Acid-base indicators and Titration

At the end of this subunit, you will be able to:

- define acid-base indicators
- write some examples of acid-base indicators
- suggest a suitable indicator for a given acid-base titration
- explain the equivalents of acids and bases
- calculate the normality of a given acidic or basic solution
- define acid-base titration
- distinguish between end point and equivalent point
- discuss the different types of titration curves.

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Acid-base indicators and Titration

Acid–Base Indicators

- Acid-base indicators are weak organic acids (denoted here as HIn) or weak organic bases (In–) that indicate whether a solution is acidic, basic or neutral.
- The color of the indicator depends on the pH of the solution to which it is added.

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- When just a small amount of indicator is added to a solution, the indicator does not affect the pH of the solution.
- Instead, the ionization equilibrium of the indicator is affected by the prevailing H_3O^+ in the solution

$$\underset{\text{Acid colour}}{HIn} + H_2O \rightleftharpoons H_3O^+ + \underset{\text{Base colour}}{In^-}$$

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- If the indicator is in a sufficiently acidic medium [increasing H_3O^+], the equilibrium, according to Le Châtelier's principle, shifts to the left and the predominant color of the indicator is that of the non-ionized form (HIn).
- On the other hand, in a basic medium [Decreasing H_3O^+] the equilibrium shifts to the right and the color of the solution will be blue due to that of the conjugate base (In^-).

$$\underset{\text{Acid colour}}{HIn} + H_2O \rightleftharpoons H_3O^+ + \underset{\text{Base colour}}{In^-}$$

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Table 1.5: Some common indicators.

Indicator Change	Acid Color	Base Color	pH Range of Color
Methyl violet	Yellow	Violet	0.0 – 1.6
Methyl orange	Red	Yellow	3.2 – 4.4
Bromcresol green	Yellow	Blue	3.8 – 5.4
Methyl red	Red	Yellow	4.8 – 6.0
Litmus	Red	Blue	5.0 – 8.0
Phenolphthalein	Colorless	Pink	8.2 – 10.0

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Example 1.10

1. What is the pH of a buffer prepared with 0.40 M CH_3COOH and 0.20 M CH_3COO^- if the K_a of CH_3COOH is 1.8 x 10^{-5} ? Which type of indicator is used to check the acidity or basicity of this solution?

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Equivalents of Acids and Bases

- **An equivalent** is the amount of a substance that reacts with an arbitrary amount (typically one mole) of another substance in a given chemical reaction.
- In a more formal definition, the equivalent is the amount of a substance needed to react with or supply one mole of hydrogen ions (H^+) in an acid–base reaction.
- For example, for Sulfuric acid (H_2SO_4) an equivalent is 2.
- For bases, it is the number of hydroxide ions (OH^-)ions provided for a reaction, for example for Barium hydroxide $Ba(OH)_2$ equivalent is equal to 2.

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- **The normality of a solution refers to the number of equivalents of solute per Liter of solution.**

$$Normality = \frac{Number\ of\ equivalents\ of\ solutes}{Liters\ of\ solution}$$

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- The definition of chemical equivalent depends on the substance or type of chemical reaction under consideration.
- Because the concept of equivalents is based on the reacting power of an element or compound, it follows that a specific number of equivalents of one substance will react with the same number of equivalents of another substance.
- When the concept of equivalents is taken into consideration, it is less likely that chemicals will be wasted as excess amount

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- Thus, according to the definition of normality, the number of equivalents is the normality multiplied by the volume of solution, in litters.
- If we add enough acid to neutralize a given volume of base, the following equation holds:

$$N_1V_1 = N_2V_2$$

- Where N1 and V1 refer to the normality and volume of the acid solution, respectively, and N2 and V2 refer to the normality and volume of the base solution, respectively.

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1. What volume of 2.0 N $NaOH$ is required to neutralize 25.0 mL of 2.70 N H_2SO_4 ?

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Acid–Base Titrations

- **An acid–base titration** is a procedure for determining the amount of acid (or base) in a solution by determining the volume of base (or acid) of known concentration that will completely react with it.
- In a titration, one of the solutions to be neutralized say, the acid is placed in a flask or beaker, together with a few drops of an acid base indicator (Figure 1.2).
- The other solution (the base) used in a titration is added from a burette and is called the **titrant**.
- The titrant is added to the acid (**Titrand** or **analyte**), first rapidly and then drop by drop, up to the equivalence point

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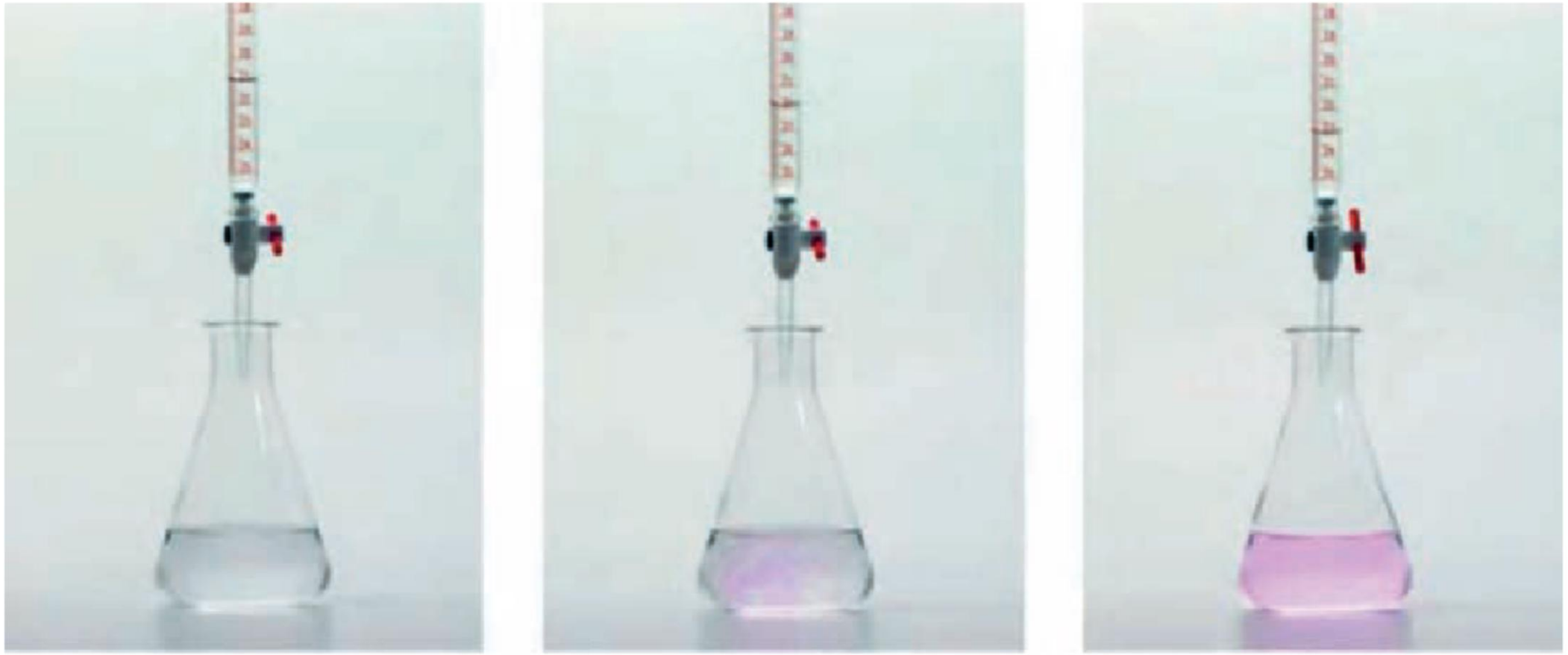


Figure 1.2: The Techniques of Titration.

- The equivalence point of the titration is the point at which the amount of titrant added is just enough to completely neutralize the analyte solution.
- At the equivalence point in an acid-base titration, moles of base are equal to moles of acid and the solution contains only salt and water.
- The equivalence point is located by noting the color change of the acid base indicator.
- The point in a titration at which the indicator changes color is called the end point of the indicator.

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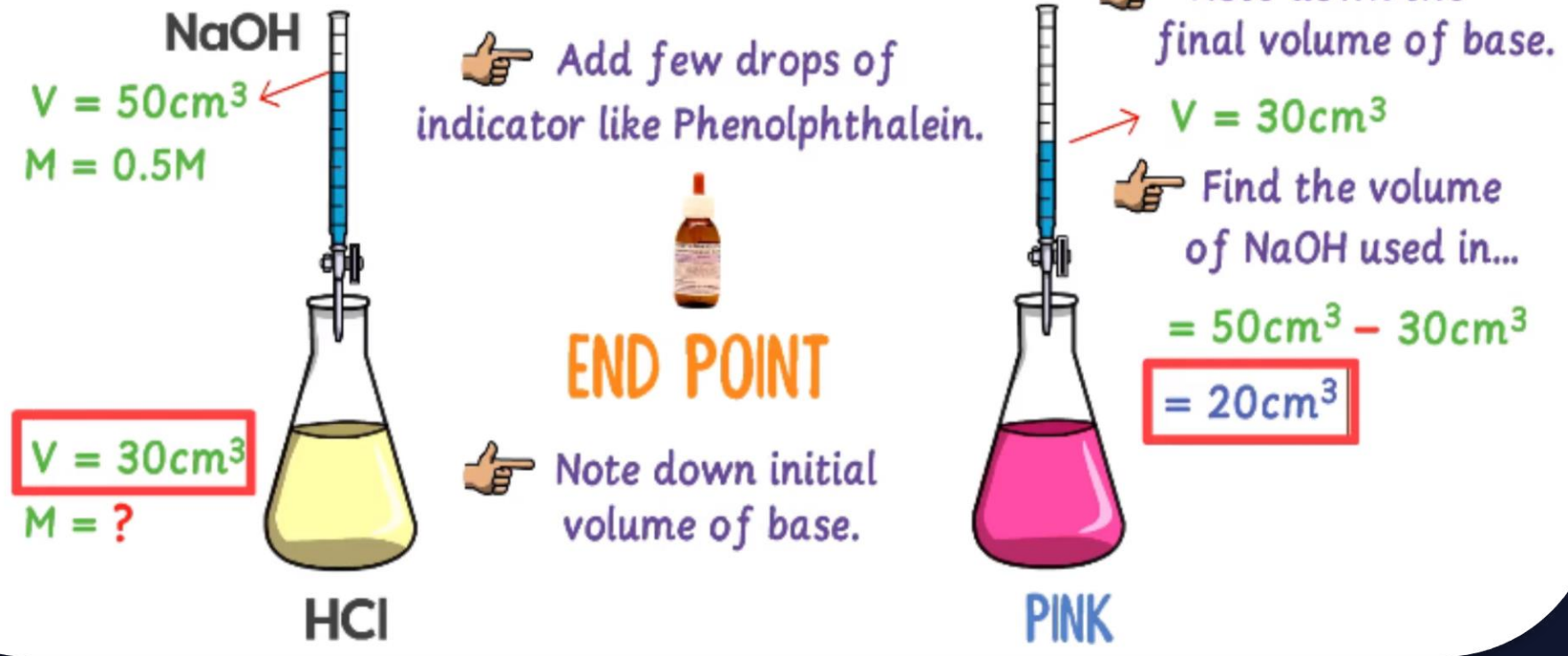
- Knowing the volume of titrant added allows the determination of the concentration of the unknown analyte using the following relation

  volume of base × concentration of base = volume of acid × unknown concentration of acid

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ACID-BASE TITRATION

A process through which the concentration of either acid or base is determined.



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UEE 2016 Question On Chemisrtry Grade 12 Unit 1

1. Which of the following is a Lewis acid?

A. SO_3^{2-}

B. SO_4^{2-}

C. BF_3

D. NH_3

UEE 2016 Question On Chemisrtry Grade 12 Unit 1

2. The pH of 0.10 M solution of an aqueous solution of a certain acid is 3. What is the value of acid ionization constant (Ka) of this acid?

A. 1.0×10^{7}

B. 1.0×10^{-5}

C. 1.0×10^{-3}

D. 1.0×10^{-1}

UEE 2016 Question On Chemisrtry Grade 12 Unit 1

3. The shift in the position of equilibrium caused by the addition of an ion already involved in the reaction is known as

A. common-ion effect
B. buffer-ion effect
C. hydrolysis effect
D. titration -effect

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UEE 2016 Question On Chemisrtry Grade 12 Unit 1

4. Which of the following is CORRECT about equivalents of acids and bases?

A. The volume of an acid or base required to reach equivalence point during acid-base titration reaction.

B. The number of moles of an acid or base required to form a one molar aqueous acidic or basic solution.

C. It is the amount of a substance that is required to react with one mole of hydroxide ions in redox reactions.

D. It is the amount of a substance that is required to react with one mole of hydrogen ions in acid - base reactions.

UEE 2016 Question On Chemisrtry Grade 12 Unit 1

5. A 250 mL solution is formed from 24.5 g of sulfuric acid (H_2SO_4). What is the normality of this solution? (Molar mass $H_2SO_4$98 g/mol)

A. 4.00 N

B. 0.250 N

C. 2.00 N

D. 0.125 N

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UEE 2016 Question On Chemisrtry Grade 12 Unit 1

6. A student prepared two beakers, each containing 100 ml of water. The student dissolved 10 mL of concentrated HCI in the first beaker and 10 ml of concentrated CHCOOH in the second beaker, which of the following would occur in the solution?

A. In the first beaker, a large fraction of HCI dissociates into ions and in the second beaker, a small fraction of CH,COOH dissociates into ions.

B. In the first beaker, a small fraction of HCI dissociates into ions and in the second beaker, a large fraction of CH COOH dissociates into ions.

C. There is no dissociation or ionization of the HCI and CH,COOH in both the first and second beakers.

D. The amount of HCI dissociated in the first beaker and the amount of CH,COOH dissociated in the second beaker are identical.

UEE 2016 Question On Chemisrtry Grade 12 Unit 1

7. What is the pOH of pure water at 37°C, where Kw equals $2.5.0 \times 10^{-14}$ and its pH is 6.8? ($\log^{2.5} = 0.4$)

A. 8.8
B. 8.2
C. 6.8
D. 7.5

UEE 2016 Question On Chemisrtry Grade 12 Unit 1

8. The hydrogen ion concentration in a certain solution at 25°C is 1.0×10^{-4} M. What is the pOH of this solution?

A. 8

B. 10

C. 2

D. 6

UEE 2016 Question On Chemisrtry Grade 12 Unit 1

9. A student collected few lemons from his/her backyard, brought it to the class and prepared a lemon juice in a beaker. They got blue and red litmus papers from his/her teacher and placed few drops of the lemon juice bots on the red and blue litmus papers, What will be their observation on the litmus papers?

A. The blue litmus paper as well as the red litmus paper will maintain their color.

B. The blue litmus paper and the red litmus paper will turn colorless

C. The red litmus paper will turn to blue and the blue litmus paper will maintain its color.

D. The blue litmus paper will turn red and the red litmus paper will maintain its color.

THE END

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