TP No. III: Acid-base titrations.

I. Introduction. Acids and bases are substances that are found everywhere and around us. In the kitchen we often use baking powder, baking soda and vinegar to make different recipes. Most of the drinks are acidic like orange juice and soda. Acids and bases are very useful in our everyday life.

II. Theoretical notions about acids and bases.

II.1. Properties of acids. They have a sour taste, their pH varies from 0 (strong acids) to 6.9 (weak acids), react with certain metals, carbonates and neutralize bases.

Examples of acids in aqueous solutions: battery acid (< 1), lemon juice (2.4 to 2.6), Coca Cola (2.5), vinegar (2.5 to 2.9), orange juice (3.5), coffee (5.0), tea (5.5), milk (6.5), pure or distilled water (7.0)

II.2. Properties of Bases. They have a bitter taste, viscous to the touch, their pH varies from 7.9 (weak bases) to 14 (strong bases), react with certain metals and neutralize acids.

Examples of bases in aqueous solutions: blood (7.34 to 7.45), sea water (8.0), soaps (9.0 to 10.0), soda (NaOH) and potash (KOH) it is 14.0

II.3. Definition of an acid and a base.

II.3.1. Definition according to Arrhenius.

An acid is a substance that can provide H^+ or (H_3O^+) ions in a solution, [acid]= [H^+]

$HCl \longrightarrow H^+ + Cl^-$

A base is a substance that dissociates in water in the form of OH^- ions, [base]= [OH^-]

$NaOH \longrightarrow Na^+ + OH^-$

II.3.2. Definition according to Brönsted-Lowry.

This theory recognizes that the acid-base reaction is a chemical equilibrium. There is a direct reaction and a reverse reaction which involves the transfer of proton (H^+). Acids are described as proton donors and bases are the acceptors of its protons.

II.3.3. conjugated acid-base couples: acid + base $\rightleftarrows$ conjugate base + conjugate acid

The conjugate base of an acid is referred to as the molecule that remains after releasing a proton. The conjugate acid of a base is the molecule that attached the proton (the proton attached to this base molecule).

Example : 1) $CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

Acid base Conjugate Base Conjugate Acid

similarly we have:

2) $CO_3^{2-}{}_{(aq)}$ + $H_2O_{(L)}$ $\rightleftarrows$ $HCO_3^-{}_{(aq)}$ + $OH^-{}_{(aq)}$

base + acid conjugate acid + conjugate base

The acid-base couple is written:

Acid/its conjugate base. Example: H_2O/ OH^-, CH_3COOH/ CH_3COO^-

Note: In the two previous examples, we notice that water sometimes acts as an acid and sometimes as a base, we say that it is an amphoteric. An amphoteric substance is a substance that acts as a proton donor (an acid) for one reaction and as a proton acceptor (a base) for another reaction. Water acts like an acid in the presence of a base and acts like a base in the presence of an acid.

II.3.4. The strength of acids and bases: A strong acid is one which in solution dissociates completely like HCl, HNO_3, $HClO_4$. A weak acid is partially dissociated like acetic acid (CH_3COOH). A strong base, in solution, dissociates completely into positive and negative ions, such as sodium hydroxide or sodium hydroxide (NaOH). A weak base provides only partial production of OH^- ions such as ammonia (NH_4OH) in solution which dissociates into two parts: NH_4^+ and OH^-. Most acids and bases dissociate very little in aqueous solution; they can be characterized by an equilibrium constant K_a (acids) and K_b (bases).

Strength acids	**strength bases**
HCl (chlorhydric acid)	NaOH (sodium hydroxide)
HBr (acid bromique)	KOH (potassium hydroxide)
HNO_3 (nitric acid)	$Ca(OH)_2$ (calcium hydroxide)
H_2SO_4 (sulfuric acid)	$Ba(OH)_2$ (barium hydroxide)

II.3.5. Acidity constant of an acid–base couple in water. It measures the strength of the acid or base. For any acid-base couple, in aqueous solution, an equilibrium is associated schematically by an equation:

AH (acid) + H_2O $\rightleftarrows$ A^- (base) +H_3O^+ ,

BOH (base)+ H_2O $\rightleftarrows$ $B^+{}_{\text{dissolved in water}}$ (base) + OH^-

The acidity constant $\mathbf{Ka = [A^-].[H_3O^+] /[AH] = K_{eq} . [H_2O]}$

II.4. Concept of pH

The pH (Hydrogen potential) scale gives a measure of acidity by measuring the concentration of hydronium ions (H_3O^+ or H^+) in an aqueous solution: pH= - log10 $[H_3O^+]$.

<u>The pH of pure water:</u> pure water has a pH equal to 7. Indeed, pure water is a neutral solution: $[H^3O^+] = [OH^-]= 10^{-7}$ moles per liter at 25 °C , pH= - $\log_{10} [H_3O^+]$= - $\log_{10}(10^{-7})$= -(-7)= +7.

Notice: * a solution is neutral when $[H_3O^+] = [OH^-]$.

** a solution is acidic when $[H_3O^+] > [OH^-]$.

*** a solution is basic when $[H_3O^+] < [OH^-]$.

II.4.1. Neutralization.

Neutralization is the dosage of a strong acid with a strong base. Both solutions have H^+ and OH^- ions that react together to form water leading to a neutral solution.

<u>The neutralization of hydrochloric acid (HCl) and sodium hydroxide (NaOH):</u>

The neutralization equation is: $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow H_2O_{(liq)} + NaCl_{(precipitated)}$

Neutralization is complete when the H^+ ions of the HCl acid react with the OH^- ions of the NaOH base, the pH obtained is equal to 7, the concentration of ions $[H_3O^+]=[OH^-]=10^{-7}M$.

II.4.2. Acid-base titration.

Titration (or dosage) is a method which makes it possible to determine or know the concentration of a substance in solution. It is an analytical method that uses a solution of known concentration (control solution) to find the unknown concentration of a substance.

In a neutralization reaction, an acid and a base are combined to form a salt and water.

The final solution is neutral. To know when the solution becomes neutral, we use an indicator (substances which, in the presence of an acid or a base, change colors) which have an acidic or basic character and which have the property of changing the color of the solution at the time of neutralization. We can also use a pH meter which will give us an exact measurement of the pH upon neutralization.

II.4.3. Determination of the concentration of the titrated solution.

To do this, we use the value of the equivalent volume, denoted $V_{equivalent}$, of the aqueous solution of the reagent found in the burette (titrant), to be poured onto a volume (test portion) of the solution of the other reagent, found in the Erlenmeyer flask (titrated), to achieve equivalence. The concentrations $N_{test\ portion}$ and $N_{burette}$ as well as the volumes $V_{test\ portion}$ and $V_{equivalent}$ are linked together by the relationship, valid at neutralization (at the equivalence point):

$$N_{acid} . V_{acid} = N_{base} . V_{base} \quad \text{or} \quad N_{test\ portion} . V_{test\ portion} = N_{burette} . V_{eq}$$

II.4.4. Colorimetric acid-base titration (use of colored indicators).

A colorimetric acid-base titration consists of identifying the equivalence using the change of an appropriate colored acid-base indicator, placed in a small quantity in the test portion of one of the reagents.

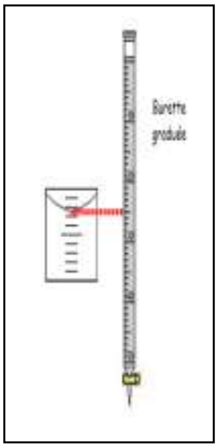
A colored indicator is suitable for titration if its turning zone contains the pH at equivalence. The jump in pH is generally very sudden at equivalence, the addition of a single drop of reagent from the burette is enough to cause the indicator to change.

II.5. Acid-base indicators.

An acid-base or color indicator is used to determine the equivalence point during a dosage.

A dye sensitive to the concentration of H_3O^+ ions tells us the pH of the solution. Acid-base indicators are usually weak acids or weak bases. The table below gives some examples of common acid-base indicators

Indicator name	pH range	color in acidic zone	color in the base zone
Methyl Orange (Helianthin)	3.1 - 4.4	Red	Yellow orange
Bromophenol blue	3.0 - 4.7	Orange-Yellow	Purple
Bromothymol blue	6.0-7.6	Yellow	Blue
Phenol red	6.4 - 8.0	Yellow	Red
Phenolphthalein (ph.ph)	8.0 -10.0	Colorless	Pink



II.6. Steps of volumetric dosage. A volumetric dosage is made up of a chain of operations:

1. determination of the concentration of the reference solution (generally known C_1),
2. filling the burette with the S_2 solution (titrant), in the burette,
3. taking a known volume V_1 of the solution S1 (titrated), place in the Erlenmeyer flask + suitable indicator,
4. titration, i.e. pouring the S_2 solution drop by drop until a change in color is observed,
5. reading V_2,

III. Operating mode.

III.1. Titration of a strong acid (HCl) with a strong base (NaOH).

1. Purpose. determine the normality and pH of hydrochloric acid (HCl).

2. Principle. It is the neutralization of a strong acid (HCl) by a strong base (NaOH).

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow H_2O_{(liq)} + NaCl_{(dissolved\ precipitate)}$

3. Operating mode.

• Place 10 mL of the HCl solution in a erlenmeyer; add 2 to 3 drops of the colored indicator (Bromothymol Blue),

• Place the NaOH solution (0.1N) in the burette using a funnel,

• Pour the soda (NaOH) slowly into the Erlenmeyer flask (drop by drop and stirring) until the color changes,

• Repeat the experiment twice.

III.2. Titration of a weak acid (CH_3COOH) with a strong base (NaOH).

1. Purpose. Determine the normality and pH of acetic acid (CH_3COOH).

2. Principle. It is the neutralization of a weak acid (CH_3COOH) by a strong base (NaOH).

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightleftharpoons H_2O_{(liq)} + CH_3COO^{-}, Na^{+}{}_{\text{(dissolved precipitate)}}$$

3. Operating mode

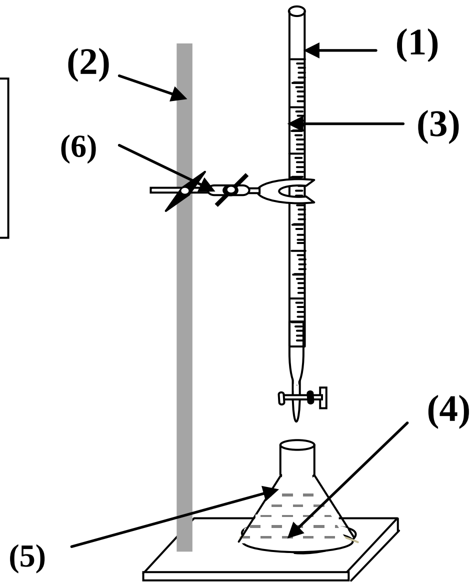
* Place 10 mL of the acetic acid solution (CH3COOH) in an Erlenmeyer flask, add 2 to 3 drops of the colored indicator (phenophthalein).

* Put the NaOH solution (0.1N) into the burette using a funnel.

* Pour the soda (NaOH) slowly into the Erlenmeyer flask (drop by drop while stirring) until the color changes

* Repeat the experiment twice.

(1) burette, (2) support (stand), (3) acid solution, (4) base solution, (5) erlenmeyer, (6)



Experimental apparatus