CHEMISTRY 203

ACID-BASE TITRATION

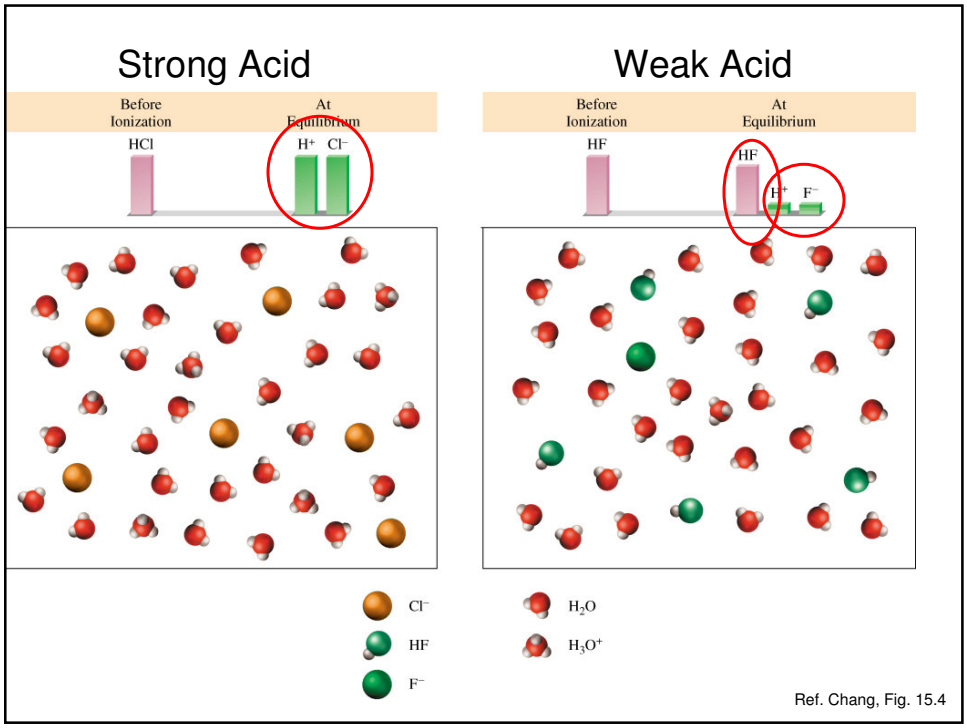
Purpose:

1. To study and practice the important analytical technique of Acid-Base titration.

2. To prepare a primary standard solution of potassium hydrogen phthalate (KHP).

3. To standardize a sodium hydroxide solution using the primary standard.

4. To determine the concentration of an unknown monoprotic acid using the standardized NaOH solution.



Ref. Chang, Fig. 15.4

Strong Electrolyte – 100% dissociation

$$NaCl\ (s) \xrightarrow{H_2O} Na^+\ (aq) + Cl^-\ (aq)$$

Weak Electrolyte – not completely dissociated

$$CH_3COOH \rightleftharpoons CH_3COO^-\ (aq) + H^+\ (aq)$$

Strong Acids are strong electrolytes

$$HCl\ (aq) + H_2O\ (l) \longrightarrow H_3O^+\ (aq) + Cl^-\ (aq)$$

Weak Acids are weak electrolytes

$$HF\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + F^-\ (aq)$$

Strong Bases are strong electrolytes

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

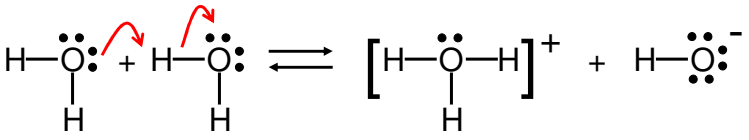
Weak Bases are weak electrolytes

$$F^-\ (aq) + H_2O\ (l) \rightleftharpoons OH^-\ (aq) + HF\ (aq)$$

Acid-Base Properties of Water

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

autoionization of water



base conjugate acid

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

acid conjugate base

Ref. Chang, Fig. 15.2

The Ion Product of Water

$$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq) \qquad K_c = \frac{[H^+][OH^-]}{[H_2O]} \qquad [H_2O] = \text{constant}$$

$$K_c[H_2O] = K_w = [H^+][OH^-]$$

The ***ion-product constant*** (***K_w***) is the product of the molar concentrations of H^+ and OH^- ions **at a particular temperature.**

At 25^0C
$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$

	Solution Is
$[H^+] = [OH^-]$	neutral
$[H^+] > [OH^-]$	acidic
$[H^+] < [OH^-]$	basic

Ref. Chang, Fig. 15.2

pH – A Measure of Acidity

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

Solution Is		At 25°C	
neutral	$[H^+] = [OH^-]$	$[H^+] = 1.0 \times 10^{-7}$	pH = 7.00
acidic	$[H^+] > [OH^-]$	$[H^+] > 1.0 \times 10^{-7}$	pH < 7.00
basic	$[H^+] < [OH^-]$	$[H^+] < 1.0 \times 10^{-7}$	pH > 7.00
	pH ↑	[H⁺] ↓	

$pOH = -\log [OH^-]$

$[H^+][OH^-] = K_w = 1.0 \times 10^{-14}$

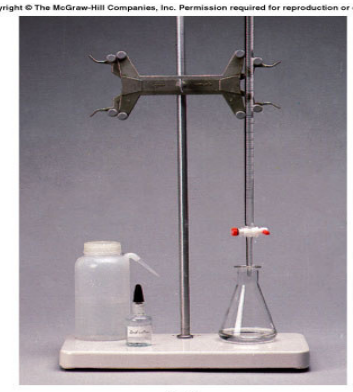
$-\log [H^+] - \log [OH^-] = 14.00$

pH + pOH = 14.00

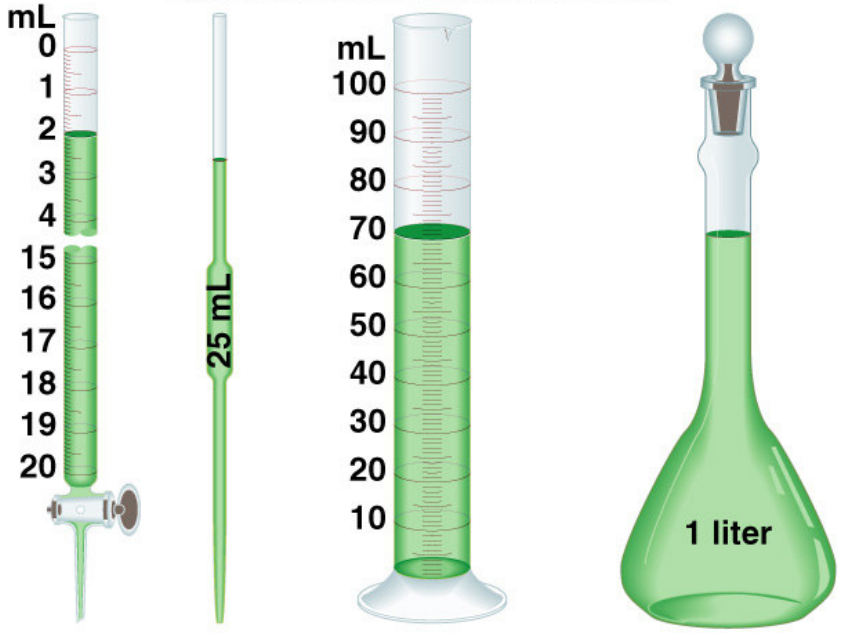
Volumetric analysis

Volumetric analysis: is a technique for determining the amount or concentration of a certain substance by doing a titration.

In titration a solution of accurately known concentration, called a primary standard, is added gradually (using a buret) to another solution of unknown concentration, until the chemical reaction between the two solutions is complete.



(a)



Equipment:

- **Mettler Balance:** Used to weigh KHP, the primary standard.
- **Volumetric Flask:** Used for making up solutions to a given volume; calibrated to contain (TC) the volume specified at the specified temperature.
- **Pipet:** Used to transfer accurately measured small volumes of solution; calibrated to deliver (TD) at the specified temperature.
- **Buret:** Used for dispensing any volume, calibrated to deliver.

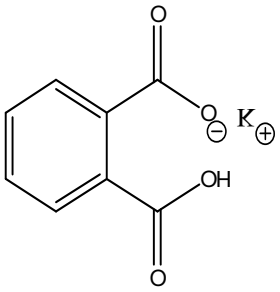
A volumetric flask, pipet, or buret is read by looking horizontally at the bottom of the meniscus of the liquid.

Standard Solution

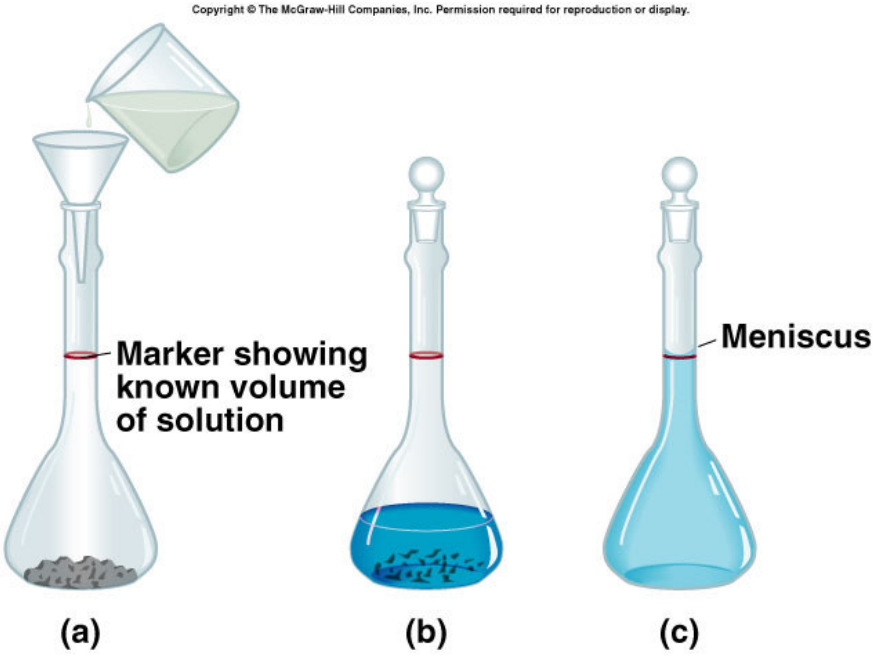
Standard solution: A solution of accurately known concentration.

Primary standard: An exact amount of a sample is dissolved in an exact volume.

e.g. Potassium Hydrogen Phthalate KHP



Molar mass = 204.23g/mol



Standard solution: (continued)

Secondary standard: Is titrated with a primary standard to determine its concentration.

e.g. Sodium Hydroxide NaOH

- Sodium hydroxide is not a good primary standard solution. Solid NaOH absorbs water from air (hygroscopic), and its solution reacts with carbon dioxide. It has to be standardized against a primary standard solution.

To standardize: To determine the exact concentration.

Properties of a good primary standard:

- High molar mass
- Stable
- Not hygroscopic
- Highly pure
- Cheap and available

Properties of a good titration:

- must be fast
- must go to completion
- should be between reactants that react stoichiometrically

Neutralization:

Acid-base titration involves a neutralization reaction which is the complete reaction between an acid and a base.

Acid + base → Salt + water

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

$KHC_8H_4O_4(aq) + NaOH(aq) \rightarrow KNaC_8H_4O_4(aq) + H_2O(l)$

or

$KHP(aq) + NaOH(aq) \rightarrow KNaP(aq) + H_2O(l)$

or

$HP^-(aq) + OH^-(aq) \rightarrow P^{2-}(aq) + H_2O(l)$

net ionic equation

When do we stop the titration?

- **Equivalence point:**

  Is the point at which all the acid has completely reacted or has been neutralized with the base, it is a theoretical concept.

- **End-point:**

  Is the point in a titration when a physical change (change in color) occurs that is associated with the condition of chemical equivalence.

Indicators:

- Indicators are used to detect the end point.
- They are substances that have distinctly different colors in acidic and basic media, they are weak organic acids or bases that change their color at the equivalence point:

$$HIn + H_2O \rightarrow H_3O^+ + In^-$$

(acidic) *(Basic)*

$$K_a = \frac{[H_3O^+][In^-]}{[HIn]}$$ K_a= Acid dissociation constant

Indicators: (continued)

- Rearranging the above equation $[H_3O^+] = Ka \frac{[HIn]}{[In^-]}$
- Taking the -log of both sides

$$- \log [H_3O^+] = - \log Ka - \log \frac{[HIn]}{[In^-]}$$

$$pH = pKa + \log \frac{[In^-]}{HIn}$$

If $\frac{[HIn]}{[In^-]} \geq 10$ the color of the acid (HIn) predominates

If $\frac{[HIn]}{[In^-]} \leq 0.1$ the color of the base (In^-) predominates

- The end point of a titration does not occur at a specific pH ; rather , there is a range within which the end point will occur, we choose the indicator whose end point lies on the steep part of the titration curve.
- In general, $pH = pKa \pm 1$

Some common acid base indicators:

Indicator	Color in acid	Color in base	pH range
Bromophenol blue	Yellow	Bluish purple	3.0-4.6
Methyl red	Red	yellow	4.2-6.3
Phenolphthalein	colorless	pink	8.3-10.0

Units and Calculations:

Molarity: *is the number of moles of solute per liter of solution*

$$M = \frac{\text{n of moles of solute}}{\text{Volume of solution in liters}}$$

$$n = \frac{Mass}{MolarMass}$$

At the equivalence point:

Complete neutralization

n of moles H^+ = n of moles OH^-

Procedure:

1. Preparation of a Primary Standard Potassium Hydrogen Phthalate (KHP) Solution:

   - Weigh a bottle containing KHP on the Mettler balance. ($KHC_8H_4O_4$, Molar mass = 204.23 g/mol)
   - Transfer the solid KHP quantitatively to a 250.0mL volumetric flask.
   - Add distilled water to 2/3 the capacity of the volumetric flask.
   - Dissolve the solid completely in the distilled water.
   - Then dilute the solution to the mark.
   - Weigh the empty bottle.

   Calculate the Molarity of KHP

2. Standardization of the NaOH Solution:

- Rinse your clean buret with distilled water and then with few ml of the NaOH solution.
- Fill the rinsed buret with NaOH, make sure there are no air bubbles.
- Record the upper buret reading.
- Pipet 25.00mLof KHP into a 250mL Erlenmeyer flask.
- Add 25 mL of distilled water.
- Add 2 drops of phenolphthalein indicator
- Titrate against KHP to the end point (light pink color).
- Record the lower buret reading.
- Repeat at least three times.

Calculate the Molarity of NaOH

3. Analysis of the unknown:

-Proceed as in part 3 using the unknown instead of KHP

Calculate the Molarity of the unknown

4. Blank Titration:

Run a blank titration to determine the blank correction for possible acids other than KHP. Proceed as in part 2, with distilled water instead of the KHP. A very small volume of the base will be needed to reach the end point. Theoretically, this volume should be subtracted from each of the proper titrations. However, since it can be as small as only one drop, you can neglect it in your calculations.