

Practical Work Chemistry II

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Objectives



This lab is intended for first-year undergraduate students in science and technology. At the end of this course, the learner will be able to:

In terms of knowledge:

- Become familiar with laboratory equipment.
- Know how to carry out a series of measurements in an organized and rigorous manner.
- Know how to use the information provided in various materials to write literal and numerical expressions, as well as equations modeling the situations proposed in the problems.

In terms of practical skills:

- Know how to use the calorimeter and the different laboratory equipment to measure the various parameters required in the problem.
- Know how to perform calculations to determine different parameters (heat capacities, specific heat capacity, concentration, pH, etc.).
- Write reports with rigor and clarity.

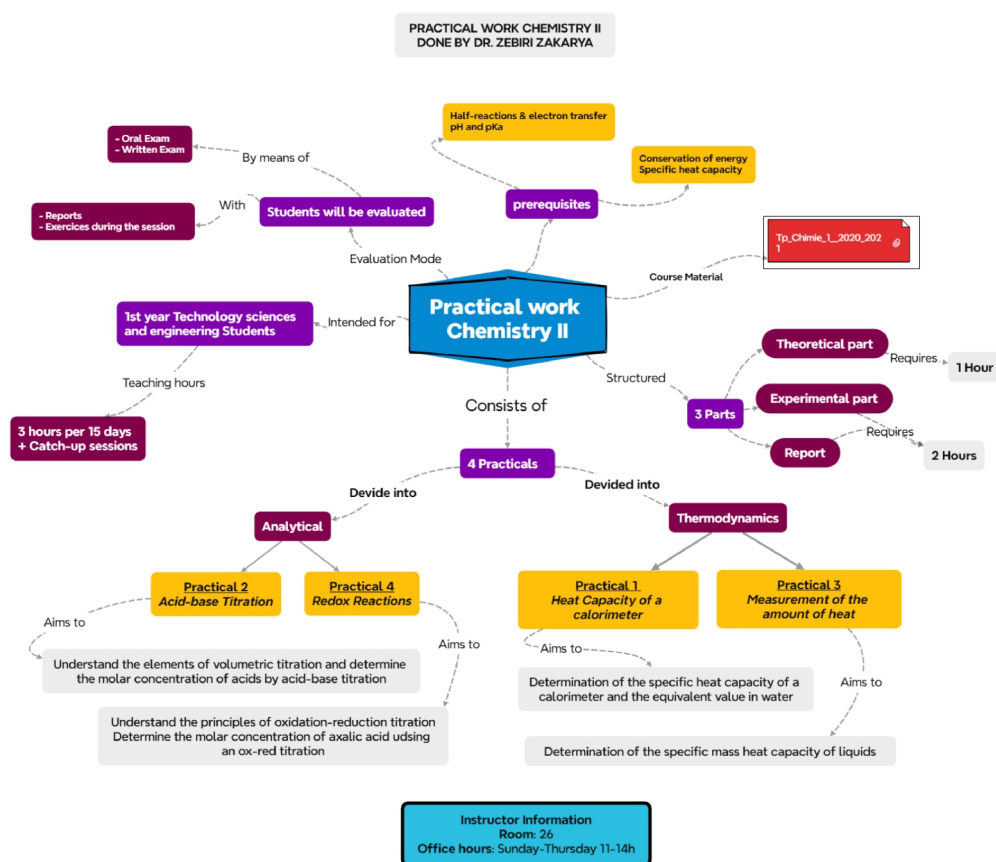
In terms of attitudes (professional behavior):

- Develop awareness of the two fundamental principles of calorimetry: the principle of equality of heat exchange (what is lost by one system is gained by another) and the principle of reverse transformations.
- Develop awareness of the principles of acid–base colorimetric titration and redox titration.

Introduction



This module introduces key experimental techniques in thermochemistry and analytical chemistry through calorimetry and acid–base titration. It covers the determination of a calorimeter’s specific heat capacity, highlighting heat transfer and energy conservation principles. It also develops skills in quantitative analysis via titration, focusing on neutralization, equivalence points, and accurate measurement.



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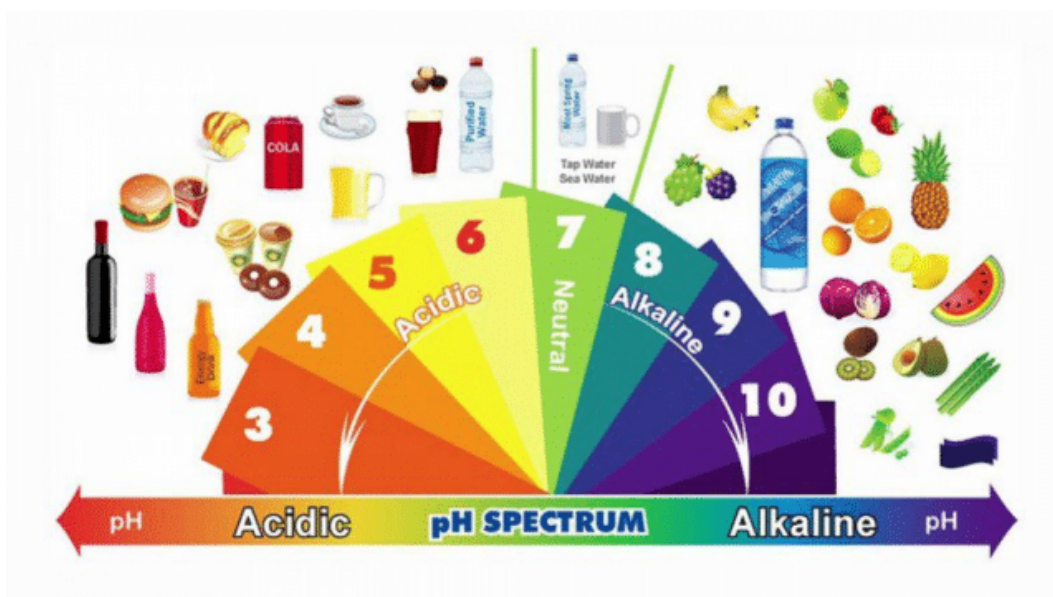
Mind map.

Practical Session No. 02: Acid–Base Titration



1. Introduction

Chemical reactions are among the most fundamental processes in the universe. A chemical reaction occurs when one or more substances, known as reactants, undergo transformation through interaction to form one or more new substances, referred to as products. Among these broad categories of reactions, acid–base reactions are particularly significant.



2. Objective of the Practical Session

At the end of this lab work, the learner will be able to:

- understand the fundamental components of a volumetric titration.
- Determine the molar concentration of an acid by means of an acid–base titration.
- Designate the appropriate colored indicator to identify the equivalence point.

3. Prerequisites

To be able to follow this practical session, learners must already know beforehand how to:

- Define acids and bases.
- Explain acid–base reaction mechanisms, as well as identify acid/base pairs.

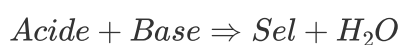
4. Quiz : Exercise

An acid is said to be when its reaction with water is complete.

5. Theoretical Section



- **Titration** is the process of determining the amount of substance, proportion, or concentration of a specific compound present in a mixture.
- **Titrimetry**, or titration, is an analytical chemistry technique used to determine the concentration of a chemical species in solution (i.e., the strength or titre of a solution). The most commonly employed method is **volumetric titration**, which involves the gradual, dropwise addition of a solution of known concentration (the *titrant*) to react with and neutralize a species present in an unknown solution (the *analyte* or titrated species).
- **Acid–base titration** is used to determine the unknown concentration of a solution containing an acid, a base, or a mixture of both. When the solution of unknown concentration is an acid, a base of known concentration is added in order to neutralize it. This type of titration is also referred to as a **neutralization reaction**, which can be represented as follows:



- The **equivalence point** can be defined as the stage in a reaction at which the number of equivalents of the titrant becomes equal to that of the analyte. It corresponds to the moment when the reaction is complete. Several methods can be used to determine the equivalence point in a titration, notably **colorimetry** and **pH-metry**. The simplest method involves the use of an indicator that undergoes a color change in the reaction mixture to signal the equivalence point (colorimetric method). The use of a suitable color indicator yields reliable results, provided that the pH range over which the indicator changes color is known.
- At equivalence, the number of moles of H_3O^+ ions provided by the acid is equal to the number of moles of OH^- ions provided by the base. This leads to the following relationship:

$$C_A \cdot V_A = C_B \cdot V_B$$

C_A : concentration of the acid

V_A : volume of the acid

C_B : concentration of the base

V_B : volume of the base



- A **color indicator** is a substance used to signal the equivalence point of a titration. Such indicators are generally selected based on their ability to undergo a visible color change at a specific pH value.
- Color indicators serve several purposes: they can be used to identify the equivalence point in an acid–base titration, and they also allow for an approximate **estimation of the pH of a solution**.
- Color indicators help determine when the equivalence point of a titration has been reached. Ideally, the indicator should change color exactly at the equivalence point. However, because pH variations occur very rapidly near this point, the addition of a single extra drop of titrant may result in a change of several pH units.

Indicator	pH range of color change	Acidic form color	Basic form color
Phenolphthalein	8.3–10	Colorless	Pink
Bromothymol blue	6–7.6	Yellow	Blue
Phenol red	6.8–8	Yellow	Red
Methyl orange	3.2–4.4	Red	Yellow

Table pH colored indicators



- The pH of a strong acid is given by: $\text{pH} = -\log C$
- The pH of a weak acid is given by: $\text{pH} = 1/2(\text{pK}_a - \log C)$

6. Experimental Section

6.1. Materials and Reagents Used

- Burette, pipette, Erlenmeyer flask, graduated cylinder, beaker
- Distilled water, color indicators, hydrochloric acid (HCl), acetic acid (CH_3COOH) ($\text{pK}_a = 4.75$), sodium hydroxide (NaOH) (0.05 M)

6.2. Experimental Procedure: Titration of a Strong Acid by a Strong Base

1. Fill the burette with the NaOH solution with a concentration of $C_B = 0.05 \text{ mol/L}$.
2. Put $V_A = 10 \text{ mL}$ of hydrochloric acid into a 100 mL Erlenmeyer flask.
3. Add 1 to 3 drops of methyl orange (MO) indicator.
4. Place the Erlenmeyer flask under the burette and add the NaOH solution while stirring manually until a color change is observed (equivalence point).
5. Record the volume of NaOH added.
6. Repeat the titration twice.

HCl		CH_3COOH	
C_A	V_B	C_A	V_B
$C1 =$	$V1 =$	$C1 =$	$V1 =$
$C2 =$	$V2 =$	$C2 =$	$V2 =$
$C_{avr} =$	$V_{avr} =$	$C_{avr} =$	$V_{avr} =$

6.3. Experimental Procedure: Titration of a Weak Acid by a Strong Base

1. Pipette 10 mL of acetic acid solution and transfer it into a 100 mL Erlenmeyer flask.
2. Add a few drops of phenolphthalein indicator.
3. Fill the burette with sodium hydroxide solution of concentration $C_B = 0.05 \text{ mol/L}$, and carry out the titration until the indicator changes color.
4. Record the volume of NaOH added.
5. Repeat the titration twice.

7. Interpretation and use of the results

1. Calculate the molar concentration C_A of hydrochloric acid and acetic acid.
2. Write the overall reaction equations for the performed titrations.
3. Calculate the pH values at the equivalence point.

8. Exercice

Quiz 1

Which region does the pH jump occur in during the titration of a weak acid by a strong base?

- ☐ from 4 to 6
- ☐ from 8 to 10
- ☐ from 10 to 12

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