République Algérienne Démocratique et Populaire
Ministère de l'Enseignement Supérieur et de la Recherche Scientifique
Université Ferhat Abbes,Setif – 1
Faculté des sciences de la nature et de la vie.
Département des études de bases.

Practical work in general and organic chemistry for L1 SNV.

Dr. Ouksel Louiza

2024-2025

L1: Natural and Life Sciences

Table of Contents

Introduction

Practical work (PW), commonly referred to as laboratory sessions or practicals, constitutes a pedagogical approach centered on experiential learning. It involves the performance of experiments designed to verify, illustrate, and complement the knowledge acquiredin theoretical courses.

Such activities are particularly relevant to the experimental sciences. In contrast to lectures or tutorials, which are conducted primarily through oral or written means, practical sessions require the use of specialized equipment (e.g., laboratory glassware, chemical reagents, or computers). The instructional space dedicated to these activities, typically designed as a laboratory, is conventionally designated as a practical work laboratory or laboratory classroom.

Practical work represents both an application and an assessment of students' mastery of the scientific method, which entails formulating hypotheses, designing experimental protocols, conducting experiments, analyzing results, and refining the initial assumptions.

Moreover, these sessions serve to demonstrate the techniques and processes implemented in applied workshops, while also allowing pedagogical materials to be adapted in accordance with the methods and technologies under investigation.

Preface

This manual of practical work is intended for first-year students in technical sciences.The main objectives of preparing and carrying out chemistry experiments are: Familiarizing the student with the most common techniques of practical chemistry Enabling the student to see how to carry out in practice some reactions already studied (or which will be) in course and/or in tutorials, therefore on a theoretical level. This should make him/her aware of the fact that practical work does not constitute a different domain from that of the course or tutorials, but rather an essential complement to the latter. This complementarity of the theoretical and practical aspects of chemistry will be reinforced by the answers to certain questions asked at the end of each experiment.

This manual covers the safety measures that must be respected in a chemistry laboratory, the different methods of preparing the solution and determining its concentration. We have tried as far as possible to limit ourselves to experiments using simple equipment easily found in educational chemistry laboratories.

Experiment 1
Safety and Practical work initiation

1.Introduction

Before discussing the various experimental techniques used during practical chemistry work, it is necessary to learn the basic safety measures to be respected during any manipulation. Safety in the laboratory is a constant problem. The chemist (student) must be aware of the potential danger of each of his actions, danger for him, for the entire laboratory and for the environment.

For the smooth running of practical work in a chemistry laboratory, it is advisable to follow the following recommendations:

1-Wearing a blouse is obligatory. The blouse must be long-sleeved, cotton and buttoned.

2-The student must know the safety measures and must know that chemicals are dangerous.

3-Eating, drinking or smoking is prohibited.

4-Do not light a flame if someone is working near flammable substances (ether, benzene, etc.).

5-Do not transfer such solvents near a flame.

6-Always turn off a burner (Bunsen burner, etc.) after use.

7-All handling with concentrated acids must be done under a hood.

8-Do not pipette concentrated acids and bases by mouth: use rubber bulbs or graduated cylinders.

9-You must arrive on time before the start of the manipulation.

10-Take your notes in a notebook and not on separate sheets.

11-Follow all the steps of the experimental part to the letter (avoiding any waste of products chemicals).

12-during the execution of the practical work, the student will note all the remarks, observations and results obtained.

13-Use only clean glassware.

14-As soon as a liquid, solution, mixture or solid is placed in a storage container, it must be immediately butcher and label it.

15-It is forbidden to throw chemicals directly into the sink.

16-All garbage (paper, matches, cotton, etc.) to be put in a trash can and not in the sink.

17-Work completed: carefully clean the bench, wash the glassware, and put the equipment back in its place.

18-At the end of the practical work, the student should be able to comment on all the results and calculations obtained (exact or inaccurate) and understand the difference between the theoretical aspect and the practical aspect.

2. Prohibitions

1. Never eat, smoke or drink in the laboratory,
2. Never pipette by mouth,
3. Never inhale a chemical product,
4. Never handle chemicals directly with your fingers or taste them.

3.Examples of common pictograms.

Safety pictograms are standardized symbols that may be encountered in laboratories or even in domestic settings, and they provide clear, essential information about the potential hazards associated with a product.

These visual warnings help users quickly identify the type of risk such as toxicity, flammability, corrosiveness, or environmental dangerwithout needing to read lengthy instructions. By offering universally recognized images, safety pictograms improve communication across languages and skill levels, ensuring that anyone handling the substance understands the necessary precautions.

Whether found on chemical containers, household cleaning products, or industrial equipment, these symbols play a critical role in preventing accidents, promoting safe handling practices, and protecting both people and the environment.

	Explosive substance. Avoid shocks and sparks!
	Oxidizing substance. This is the case for many oxidants such as oxygen; they react with combustible materials.
	Flammable substance. This is the case for many organic compounds (alkanes, alcohols, etc.). They must be kept away from sources of heat (flame, spark).
	Toxic substance. Avoid any contact with skin and eyes; do not inhale the vapors (work must be carried out under a fume hood).
	Irritant or harmful substance. Same precautions as for toxic products; work in a well-ventilated area.
	Corrosive substance. This is the case for concentrated acids and bases. Avoid any contact with skin and eyes. Wear a lab coat and use gloves.

4.Practical work initiation.

Enumerate the various types of laboratory glassware and describe the purpose of each.

Beaker:A simple container used for mixing, stirring, and heating liquids.

Erlenmeyer flask (conical flask):Used for mixing liquids by swirling, heating, and carrying out titrations.

Volumetric flask:Designed to prepare solutions of precise volumes and concentrations.

Graduated cylinder:Used to measure the volume of liquids accurately.

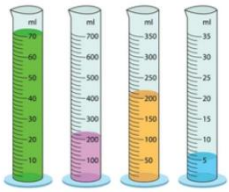
Burette:Used to dispense precise volumes of liquid, especially in titrations.

Pipette (volumetric or graduated): Used to transfer or measure precise volumes of liquid.

Funnel: Used for pouring liquids into containers with small openings or for filtering with filter paper.

Watch glass: Used to hold small samples, cover beakers, or evaporate liquids.

There are also other laboratory accessories, such as: a propipette, a wash bottle, test tubes, etc...

			
Volumetric flask	Beaker	Erlenmeyer flask	Graduated cylinder
			
Pipette	Burette	Funnel	Watch glass

5. Guided tour

-Together with the teacher, identify the characteristics on the label of a chemical product.

-Turn on the fume hood and explain how it works.

-Fill with distilled water and explain the use and reading of volumes on different types of glassware while respecting precision.

6.Preparation of solutions

-Define the concepts of molarity, normality, molality, and mass fraction. Provide examples of calculations and unit conversions.

-Determine the mass of NaCl that must be weighed to prepare a solution with a concentration of 0.5 $mol \cdot L^{-1}$ in a 100 mL volumetric flask ($M_{NaCl} = 58.5\ g \cdot mol^{-1}$). Take into account the purity of the table salt used.

-Prepare a solution with a concentration of 0.25 $mol \cdot L^{-1}$ from the 0.5 M stock solution.

-Prepare a solution with a concentration of 0.01 $mol \cdot L^{-1}$ of acetic acid.

7.Report

-List, in order, all the equipment used for the preparation of the NaCl stock solution and the diluted solution.

-What volumes of the 0.5 M NaCl stock solution should be taken to prepare solutions with concentrations of 0.1 M and 0.2 M, for final volumes of 50 mL and 100 mL?

-Give all the steps followed for the preparation of the acetic acid solution with the minimum possible errors.

-What is the value of the error committed in the calculation of concentrations, given the absolute uncertainties of the glassware used:

Equipments	Absolute uncertainty
Volumetric flask (250 mL)	$\Delta V = 0.15mL$
Volumetric flask(100 mL)	$\Delta V = 0.10mL$
Pipette (5 mL)	$\Delta V = 0.02mL$
Pipette (10 mL)	$\Delta V = 0.05mL$
balance	$\Delta m = 0.1\ g$

Knowing also that the dilution equation corresponds to the equality of the number of moles of the compound before and after dilution, as follows:

$$C_i V_i = C_f V_f$$

$$C_f = C_i . V_i / V_f$$

By introducing the logarithmic function and then the associated differentials.

Experiment 2
Geometric by molecular models

1.Introduction

The primary aim of this laboratory work is to construct molecular models using ball-and-stick representations in order to visualize the three-dimensional arrangement of atoms within molecules. Such models provide valuable insights into the study of stereoisomerism, particularly optical and geometrical isomerism. By examining the spatial orientation of atoms, this approach highlights the crucial role of molecular geometry in determining the physical and chemical properties of compounds.

2.Objective

Build simple molecules using molecular models. Determine, from its electronic structure, the bonds engaged by an atom within a molecule. Study the spatial geometry of molecules in relation to these atomic bonds.

3.Materials

A box of molecular models with atoms modeled by balls with conventional colors:

Atom	Hydrogen	Carbon	nitrogen	Oxygen
Color	White	Black	Blue	Red

4.Principle

It is possible to construct an exploded model: it allows you to clearly visualize the bond angles.

5.Operating mode

The main objective of this practical work is to make molecular models using rods and small balls and to study the different types of geometry.

5.1. Carbon C central atom

- Give the Lewis models of the following molecules: CH_4 Methane, C_2H_6 Ethane, C_2H_4 Ethylene, C_2H_2 Acetylene.

-Construct the corresponding exploded models.

-Around each Carbon, give the name of the spatial geometry of the bonds.

5.2.Nitrogen N central atom

-Make the molecules: NH_3 Ammonia, CH_3 -NH_2 Methanation.

-Around each Nitrogen, give the name of the spatial geometry of the connections.

5.3. Oxygen O central atom

- Make the following molecules: H_2O Water, CO_2 Carbon Dioxide.

-For these models: - What is the geometry of the Water H_2O molecule?

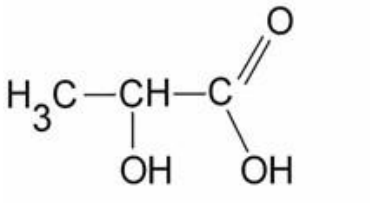
- What is the geometry of CO_2?

Experiment 3
Milk Acidity -Test for Quality and Freshness

1.Introduction

Milk is widely produced and consumed worldwide as dairy products. Milk has many health benefits, as it is rich in protein, vitamin D and calcium, which are good for bone health and development.

During the microbial multiplication in milk the lactose would be converted into lactic acid which would result in increase in the acidity of milk and decrease in the pH value. This acidity is known as developed acidity. The normal acidity of individual cow Milk ranges from
0.10 to 0.26% lactic acid. Milk having titratable acidity more than 0.18% Lactic acid is not suitable to prepare heat treated products as the milk will coagulate at or above that acidity.



Lactic acid or 2-hydroxypropanoic acid.

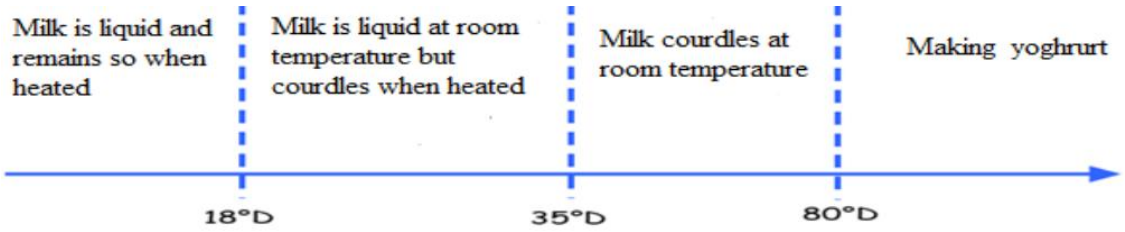
2.Objective

The aim of this experiment is to determine the concentration of lactic acid in a commercial milk sample by titration, that is, the gradual addition of a standard titrant solution contained in a burette, to the analyte solution of unknown concentration until all the reaction is complete.

2.1. Dornic Acidity

One of the most common ways to express the acidity of dairy products is in Dornic degrees (°D). One °D is equal to 0.1 g of lactic acid per liter of milk. The acidity of good quality fresh is usually below 18°D.

$$D° = \frac{Cm}{0.1}$$



2.2. Titration of the milk sample

- Fill the burette with 0.1 M NaOH solution.

- Use a 10 mL pipet to transfer 10 mL of milk to a beaker. Use the same pipet, without rinsing it first, to add 90 mL of distilled water to the milk sample.
- Add 5 drops of phenolphthalein and put a white paper beneath the beaker.

- Add NaOH dropwise from the burette, swirling the beaker, until the color of the solution in the beaker changes to pink. The equivalence point has been reached when the pink color does not fade for at least 30s.

3. Results and calculations

1. Write the reaction involved in the titration.
2. End deduce the dilution factor.
3. Calculate the molar concentration of lactic acid in the milk after and before the dilution (use the following formula:

$$C_a \times V_a = C_b \times V_b).$$

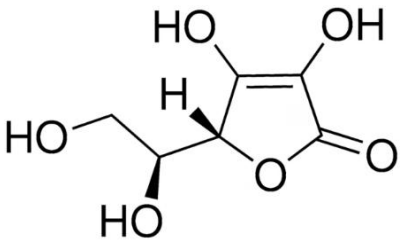
4. Calculate the mass concentration of lactic acid and then the Dornic acidity of the milk.
5. Mark the position of the milk tested on the proposed scale with a cross. And comment.

Data: Molar mass of H = 1g/mol, C = 12 g/mol, O = 16 g/mol and Na = 23g/mol.

Experiment Vitamin C dosage	4

1. Introduction

Vitamin C is an organic molecule also called ascorbic acid with the formula:



Acide ascorbique.

It is an antioxidant: it prevents dioxygen from oxidizing food and therefore altering it. Its presence in food is indicated by the code **E300**. Vitamin C behaves in water as a weak acid. This property can be used to determine the mass of vitamin C contained in a tablet.

2.Objective

The objective of this manipulation is to analyze the vitamin C contained in a tablet and to calculate the percentage of ascorbic acid present in this pharmaceutical preparation.

3.Experimental

3.1. Materials

-Pestle and mortar (glass rods can be used).

-500 mg vitamin C tablets.

-Funnel and 100 ml graduated flask.

-Distilled water.

-Graduated burette.

-Sodium hydroxide solution (NaOH) with molar concentration C_b = 0.05 mol/l.

-Pipette.

-Beaker or Erlenmeyer flask.

-Phenolphthalein.

-Magnetic stirrer + magnet.

-Balance.

3.2.Operating mode

- Weigh a vitamin C tablet.
- Crush the tablet into powder using a pestle and mortar or a glass rod.
- Place the powder in a 100 ml graduated flask, add distilled water to 3/4 of the flask and shake.
- Top up to the mark with distilled water
- Take a volume VA = 10 ml of the vitamin C solution and pour it into an Erlenmeyer flask.
- Add 2 to 3 drops of phenolphthalein.
- Fill the burette with the titrant solution (sodium hydroxide) of molar concentration CB = 0.05M. Making sure that there is no air in the lower part.
- Adjust the liquid to the zero level of the burette by draining the excess reagent into a beaker.
- Add the titrant solution ml by ml until the indicator changes color.

(Colorless to pink).

At equivalence we have :

$$\mathbf{C_A * V_A = C_B * V_B}$$

The molar mass of vitamin C is given : $M(C_6H_8O_6) = 6*12+8+6*16 = 176$ g/mol.

4. Questions

1. Write the reaction between ascorbic acid and sodium hydroxide (use the chemical formula for ascorbic acid).

2. Calculate the mass concentration of the vitamin C solution (before dosing).

3. Determine the amount of matter present in the vitamin C (before dosing).

4. Calculate the molarity of vitamin C (after dosing).

5. Calculate the mass concentration of ascorbic acid present in the tablet (after dosing).

6. Deduce the mass of ascorbic acid present in a tablet.

7. Determine the percentage by mass of ascorbic acid in a tablet.

8. Compare and comment on the result obtained with that written on the vitamin C label.

Experiment 5
Titration of strong acid with strong base

1.Introduction

Titration describes a process where the concentration of an unknown substance is determined by comparing it with a solution of known concentration. The concept that makes titrations possible is finding the equivalence point, i.e., identifying when the quantity of the unknown substance is equal to the quantity of the known substance.

The equivalence point is found in a titration by adding trace amounts of a substance, called an indicator, which turns color when the equivalence point is reached. When a strong acid is titrated with a strong base, or vice versa, the pH of the solution will be about 7.0 at the equivalence point. Phenolphthalein is the indicator used in this experiment. Phenolphthalein is colorless in acidic solutions and turns pink in alkaline solutions.

2.Objectives

-Determine the molarity of an unknown hydrochloric acid solution.

-Understand the use of indicators in titrations.

3.Principal

This experiment will be done in two parts: (1) Preparation of a 0.1 M sodium hydroxide solution by dilution and (2) determination of the Hydrochloric Acid concentration in an unknown solution by titrating it with a solution of known concentration.

Acids: contain hydrogen and dissociate in water to give H^+:

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

Bases: contain the hydroxyl group and dissociate in water to give OH^- :

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

A Strong acid (such as HCl) and a Strong base (such as NaOH) react to form water

and a salt. In this reaction, we say the acid and the base are **neutralized**:

$$HCl(aq) + NaOH(aq) \rightarrow Na^{+}(aq) + Cl^{-}(aq) + H_2O$$

4.Experimental

4.1.Materials

Beaker or Erlenmeyer flask, Funnel, Buret, Volumetric Pipette, Distilled Water, Phenolphthalein ($C_{12}H_{14}O_4$), Sodium hydroxide (NaOH) and Hydrochloric acid (HCl) solutions.

4.2. Procedure

Part I: Preparation of 0.1 M Sodium Hydroxide Solution.

Prepare 50 ml of a 0.1M from prepared 1M NaOH.

Part II: Titration of Hydrochloric acid

1. Rinse your burette and fill it with 0.1 M NaOH solution. (Note: be sure there are no air bubbles in the tip of the burette).

2. Obtain a 10 mL pipet and transfer 10 mL of HCl solution into the Erlenmeyer flask.

3. Add 1-3 drops of phenolphthalein indicator solution into the Erlenmeyer and place it under burette.

4. Begin the titration, swirling the solution in the Erlenmeyer flask as you add NaOH in a drop-wise approach (Note: The solution will initially turn pink, and then fade back to colorless when swirled. The pink color will remain longer as you approach the end point of the titration).

5.Record the final volume of NaOH in the burette when the solution in the Erlenmeyer flask remains light pink.

Calculations:

Use the following relationship to calculate the Hydrochloric acid concentration:

$$\mathbf{V_a x\ C_a = V_b\ x\ C_b}$$

Molar mass of HCl = 36.4g/mol and NaOH = 40g/mol

5.Report

1.What mass of the 1M NaOH solution must be weighed to prepare 1L solution?

2.Calculate the mass percent of the solution.

3. En deduce the dilution factor to prepare 0.1M NaOH from 1M solution. 4.Calculate the molarity of HCl.

5.Calculate the number of moles of Hydrochloric acid. 6.Calculate the Concentration in grams per liter.

Experiment 6

Determination of acidity in vinegar

1.Introduction

A common household item, vinegar, contains acetic-acid in addition to other chemicals. The purpose of this experiment is to identify the amount and mass percent of acetic acid (CH_3COOH) present in vinegar. This will be accomplished by titrating the solution with a standard base, Sodium hydroxide. In this experiment, a technique known as a titration will be used to determine the concentration of acetic acid in vinegar. A titration involves performing a controlled reaction between a solution of known concentration (the titrant) and a solution of unknown concentration (the analyte). Here, the titrant is an aqueous solution of 0.1 M sodium hydroxide (NaOH) and the analyte is vinegar. When mixed, a neutralization reaction occurs between sodium hydroxide and the acetic acid in vinegar:

$$CH_3COOH + (Na^+ + OH^-) \longrightarrow CH_3COO^- (Na^+) + H_2O$$

The NaOH will be added to the vinegar sample until all the acetic acid in the vinegar has been exactly consumed. At this point the reaction is completed, and no more NaOH is required. This is called the equivalence point of the titration.

In order to know when the equivalence point is reached, an indicator solution called phenolphthalein is added to the vinegar at the beginning of the titration. Phenolphthalein is colorless in acidic solutions like vinegar, and deep pink in basic solutions like sodium hydroxide. At the equivalence point of the titration, just one drop of NaOH will cause the entire solution in the beaker to change from colorless to a very pale pink.

2.Experimental

2.1.Materials

Burette, beaker, graduated pipette, funnel, 0.1 M NaOH solution, vinegar, phenolphthalein and wash bottle with distilled water.

2.2.Procedure

-Fill the burette with NaOH solution.

-Use the graduated pipette to transfer 10 mL of vinegar into a clean beaker. Then add 20 mL of distilled water and 3 to 5 drops of phenolphthalein to this beaker.

-Begin the titration by slowly adding NaOH from the burette to the vinegar in beaker. Swirl beaker as you add the base in order to efficiently mix the chemicals. Some pinkness may appear briefly in the beaker as the base is added, but it will quickly disappear as the beaker is swirled. As the equivalence point is approached, the pink color will become more pervasive and will take longer to disappear. When this occurs, start to add the NaOH drop by drop. Eventually the addition of just one drop of NaOH will turn the solution in the beaker a pale pink color that does not disappear when swirled. This indicates that the equivalence point has been reached. Do not add any more NaOH solution at this point. Record this volume of NaOH solution.

3.Questions

1. Write the reaction involved in the titration.

2. What was the purpose of the phenolphthalein indicator in this experiment?

3. Suppose you added 40 mL of water to your vinegar sample instead of 20 mL. Would the titration have required more, less or the same amount of NaOH for a complete reaction?

4. Calculate number of moles of NaOH used in titration.

5. Calculate molarity of CH_3COOH in vinegar (use the following relationship to calculate the concentration: $\mathbf{C_a\ xV_a = C_b\ x\ V_b}$).

6. Calculate mass of acetic acid.

7. Deduce mass percent of CH_3COOH in vinegar.

- Molar mass of H = 1g/mol, C = 12 g/mol, O = 16 g/mol and Na = 23g/mol.
- Density of vinegar = 1.00 g/mL.

Experiment 7
Redox-Titration
(determination of hydrogen peroxide concentration)

1.Introduction

Hydrogen peroxide (H_2O_2) is one of the most powerful oxidants used in many sectors such as clean wounds. Because it readily decomposes in the presence of light, heat, or metallic catalysts into water and oxygen, the quantity of a hydrogen peroxide solution must be checked regularly to maintain its effectiveness. The concentration of hydrogen peroxide can be analyzed by redox titration with potassium permanganate.

Oxidation-reduction (redox) reactions involve a transfer of electrons between the species being oxidized and the species being reduced. The reactions are often balanced by separating the reaction components into half-reactions: oxidation (loss of electrons) and reduction (gain of electrons). In a redox reaction, the number of electrons lost by a species being oxidized is always equal to the number of electrons gained by the species being reduced. In the reaction being studied in this lab, solutions of hydrogen peroxide, H_2O_2, and potassium permanganate, $KMnO_4$, will be combined in acidic solution.

The $KMnO_4$ solution is purple in color, as it is added to the hydrogen peroxide solution, it reacts with the H_2O_2 to produce colorless Mn^{2+}, and thus the color will fade in the flask. When all the H_2O_2 has been used up, the last drop of potassium permanganate that is added will keep its color. The endpoint of the titration is the point at which the last drop of $KMnO_4$ added to the solution causes it to turn pink.

At equivalence, the number of moles of permanganate ions and hydrogen peroxide present are in the stoichiometric proportions given in the balance equation.

$$C_1 = \frac{5 \text{ x } C_2 \text{ x } Ve}{2 \text{ x } V_1}$$

Were:

C_1: molar concentration of hydrogen peroxide V_1: Volume of hydrogen peroxide

C_2: molar concentration of permanganate Ve: equivalent volume.

2.Objectives

Analytical method such as titration by permanganometry is used to maintain the quality of products containing H_2O_2. Thus, the objective of this practical work is to evaluate commercial sample of H_2O_2 10V by permanganometry.

3.Experimental

3.1. Materials

-Hydrogen peroxide solution (sample to be analyzed).

-Potassium permanganate ($KMnO_4$) solution.

-Sulfuric acid (H_2SO_4), dilute (usually 1–2 M).

-Distilled or deionized water.

-Burette (50 mL).

-Burette stand and clamp

-Pipette (10 or 20 mL) .

-Erlenmeyer flask (250 mL).

-Volumetric flask (for preparing $KMnO_4$ solution, if needed).

-Beaker(s) (100–250 mL).

-Funnel.

-Pasteur pipette.

3.2.Procedure

- Fill buret with 0.02 M $KMnO_4$. Be sure to get rid of any air bubbles by opening the stopcock.
- Place 1 mL of the hydrogen peroxide solution into Erlenmeyer flask.
- Add about 25 mL of distilled water to the flask.
- Place 5 mL of 6 M H_2SO_4 into the flask.
- Swirl the contents of the flask.
- Titrate the hydrogen peroxide solution with the $KMnO_4$ solution. Record the final

volume of $KMnO_4$.

4.Report

1- Show the two redox couples involved and write the associated redox half-equations.

2- Balance the following redox reaction in acidic solution.

$$MnO_4^-(aq) + H_2O_2(aq) \rightarrow Mn^{2+}(aq) + O_2(g)$$

3- How many electrons are transferred in this reaction?

4- What species is being oxidized? What species is being reduced?

5- Calculate molarity and normality of hydrogen peroixde.

6- Hydrogen peroixde decomposes on heating according to the equation:

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

Determine the volume of oxygen (O_2) that can be supplied by 1 L of the commercial solution. And compare with the bottle of drugstore.

Data: Molar volume of gases Vm = 22.4 L. mol^{-1}

Experiment 8
Solubility (Determination of chloride ions)

1. Introduction

It is well known that chlorides have effects on the human body. Sodium chloride is responsible for certain kidney problems, cervical cancer, and most notably, hypertension.Solubility is the maximum amount of a salt that can be dissolved in one liter of water. Beyond this amount, precipitation occurs. A solution is said to be saturated when it is at the limit between precipitation and solubility. A slightly soluble salt is characterized by a solubility product, denoted as **Ks**. The expression of Ks depends on the reaction involved. Let the reaction be:

$$MnXm \leftrightarrow \quad nM^{m+} + \quad mX^{n-}$$

2. Objective

The objective is to determine the concentration of chloride ions (Cl^-) contained in mineral water and tap water using a solution of Ag^+ through the formation of a slightly soluble salt. Calling the concentration formed "s" solubility, The solubility product Ks is expressed as:

$$Ks = [M^{m+}]^n.[X^{n-}]^m = (nS)^n.(mS)^m = n^n.m^m.S^{n+m}$$

For example, for the salt Ag_2CrO_4 we have:

$$Ks = [Ag^+]^2.[CrO_4^{-2}] = (2s)^2.s = 4s^3 = 10^{-12}$$

From which we can derive the value of solubility s from Ks.
For the salt Ag_2CrO_4, for example:

$$S = \sqrt[3]{\frac{K_s}{4}}$$

In the case of AgCl, we write:

$$Ks = [Ag^+].[Cl^-] = s^2 = 10^{-10} \Rightarrow S = \sqrt{K_s}$$

The solubility of a salt in water depends on temperature, pH, and the presence of a common ion. The latter generally decreases solubility.The opposite phenomenon, that is precipitation, also depends on the concentration of ions. Precipitation occurs

when the ionic product (Pi) is greater than Ks. On the other hand, solubility occurs when Pi is less than Ks. At equality,the solution becomes saturated. By comparing the solubility values of the two salts given as examples, it is observed that chloride ions (Cl^-) precipitate first before chromate ions (CrO_4^{2-}) in a solution containing Ag^+.

$$Pi = [M^{m+}]^n.[X^{n-}]^m = Ks$$

3.Experimental Part

In the following, we propose to determine chloride ions using two different methods: The first method, **Mohr's method**, with a silver nitrate ($AgNO_3$) solution of known concentration.The second method consists of titrating chlorides with a silver nitrate solution of unknown concentration, using a standard potassium chloride (KCl) solution of known concentration.

3.1.Mohr's method

3.1.1.Required material

Glassworks	Products
Burette	K_2CrO_4 0.1M
Erlenmeye	$AgNO_3$ 0.1M
Pipette	Mineral water
Pissette	Tap water
Test tube	

Nitrates are completely soluble, and the $AgNO_3$ salt then dissociates completely to form Ag^+ and NO_3^-. The Ag^+ ion combines with Cl^- to precipitate as the AgCl salt.The presence of CrO_4^{2-} is only detectable after all the chlorides have been consumed. At this point, an additional amount of Ag^+ is added, leading to the precipitation of the Ag_2CrO_4 salt, characterized by its color. We can then write the two successive equations:

$$Ag^+ + Cl^- \rightarrow AgCl_{(s)} \text{ (White precipitate)}$$

$$2Ag^+ + CrO_4^{-2} \rightarrow Ag_2CrO_{4(s)} \text{ (Red precipitate)}$$

3.1.2.Experimental protocol

-Prepare in a 50 mL flask a diluted 0.01 M $AgNO_3$ solution from the 0.1 M stock solution.

-Fill the burette with the 0.01 M silver nitrate solution.Pour 50 mL of mineral water to be analyzed into an Erlenmeyer flask.

-Add 1 mL of potassium chromate to the Erlenmeyer flask.

-Add the silver nitrate solution drop by drop until the color changes.

-Record the volume at the equivalence point.

It should be noted that the detection of the color change at the equivalence point can be improved by preparing a reference solution in a beaker by mixing 1 mL of 0.1 M $AgNO_3$, 1 mL of 0.1 M K_2CrO_4, and 20 mL of mineral water.

- Dose 50 mL of tap water in the same manner. Compare the results. At the end of the measurements, rinse the burette and all glassware thoroughly with distilled water.

3.2. Titration of Chlorides Using a Silver Nitrate Solution of Unknown Concentration

3.2.1.Required material

-A graduated burette, an Erlenmeyer flask.

-Distilled water.

-Potassium chloride (KCl) solution at 150 $mg \cdot L^{-1}$

-Silver nitrate ($AgNO_3$) solution of unknown concentration.

-Potassium chromate. (K_2CrO_4 0.1 M)

3.2.2.Experimental protocol

a)Standardization of the silver nitrate solution:

-Introduce 50 mL of the 150 $mg \cdot L^{-1}$ KCl solution into an Erlenmeyer flask.

-Add 4 drops of potassium chromate to the Erlenmeyer flask.

-Fill the burette with the silver nitrate solution.

-Slowly add the silver nitrate solution to the Erlenmeyer flask while gently swirling by rotational movements.

-Immediately close the burette tap when the color indicator changes to a brick-red shade.

-Record the volume $\mathbf{V_1}$ of silver nitrate required to reach this endpoint.

b)Titration of Chlorides in the Water Sample to Be Tested:

-Repeat the previous procedure, replacing the 50 mL of KCl solution with 50 mL of the water sample to be analyzed (mineral water or tap water).

-Proceed in the same way as before.

-Record the volume $\mathbf{V_2}$ of silver nitrate required to reach the endpoint.

-Collect the excess silver nitrate and rinse the burette with distilled water.

c) Data Analysis:

Knowing the concentration of the KCl solution and the volume $\mathbf{V_1}$ of silver nitrate required to reach the endpoint, calculate the chloride ion concentration in the analyzed water sample by using the measured volume $\mathbf{V_2}$.

4. Report

- Silver nitrate is a solid. How much silver nitrate is needed to prepare 500 mL of a solution with a concentration of 0.1 mol.L-1?

- In practice, for water determination, a 4.8 g.L-1 silver nitrate solution is used. The volume, in mL, of the silver nitrate solution required for the determination then directly gives the number of mg of Cl- ions determined in the sample. Justify this method with calculations. Given molar mass (g/mol): Ag = 108, Cl = 35.5, N = 14.0, O = 16.0

- Calculate the respective solubilities of the salts AgCl and Ag_2CrO_4.

- Record the volumes of $AgNO_3$ added for the mineral water and tap water. Calculate the Cl^- concentration and compare the two methods.

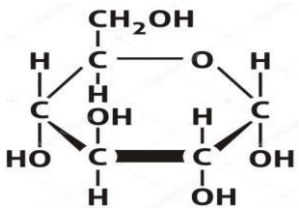
- Express the chloride content in $mg.L^{-1}$.

Experiment 9
Determination of glucose in solution by spectrophotometry

1.Introduction

Glucose is an energy molecule with the chemical formula **$C_6H_{12}O_6$**. It is used as a source of energy by most living beings, through cellular respiration. This is then completely oxidized into H_2O and CO_2.

WM $_{C_6H_{12}O_6}$ = 180 g/mol



Glucose

Blood glucose level is a value commonly measured during a blood test. It is regulated by various physiological mechanisms, notably by insulin. The result depends on whether being fasted or having eaten shortly before a blood test has an influence on the result;

Elevated fasting blood sugar may be a sign of a prediabetic state or diabetes mellitus.

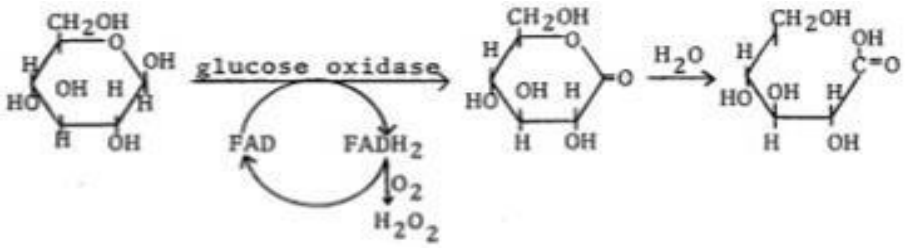
Conversely, blood sugar levels that are too low are called hypoglycemia.

2.Objective

Determine the concentration of a glucose solution present in a 5% glucose serum (quality control).

3.Principle of GOD-POD

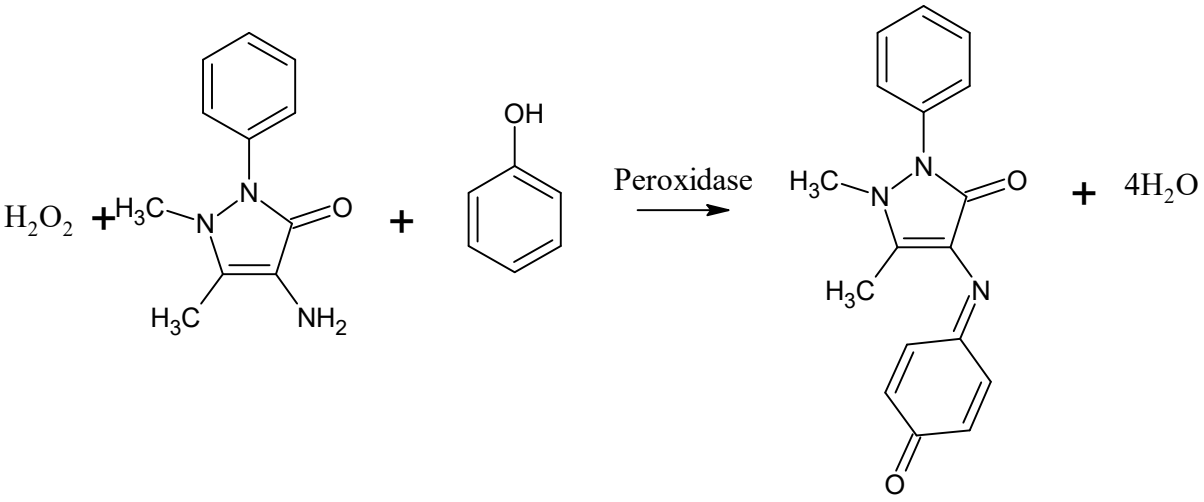
In the 1st reaction: Glucose oxidase (GOD) catalyzes the oxidation of glucose into gluconic acid.



Glucose-oxidase (GOD)

Glucose + O_2 ⟶ acid gluconic+ H_2O_2

In the 2nd reaction: The hydrogen peroxide (H_2O_2) produced in the 1st reaction is detected with a chromogenic oxygen acceptor, phenol, 4–aminophenazone (4-Aminoantipyrine), in the presence of peroxidase (POD) which produces Quinoneimine (which absorbs in λ= 505 nm) + H_2O:



$$\text{4-Aminoantipyrine} + 2H_2O_2 + \text{phenol} \xrightarrow[\text{(POD)}]{\text{Peroxidase}} \text{Quinoneimine } (\lambda = 505\text{ nm}) + 4H_2O$$

Rose — Uncolored

The intensity of the pink color (quinonimine) produced is proportional to the concentration of glucose present in the sample tested, as measured at λ= 505 nm.

Materials and reagents: Usual laboratory equipment: Water bath; UV-visible spectrophotometer plus a cuvette (an optical path of 1 cm); adjustable pipettes and/or different volumes (from 20 to 1000 µl), Blue and yellow tips, test tubes and racks; marker dosing and calibration reagents R1 and R2 and distilled water.

Composition of reagents (ready to use):

Reagent1: R1: TRISa buffer: 92 mmol/L, pH 7.4; Phenol: 0.3 mmol/L. Enzymes: Glucose Oxidase (GOD): 15000UI/L, Peroxidase (POD): 1000UI/L; 4-Aminophenazone: 2.6 mmol/L. (a) TRIS: Tris(hydroxymethyl)aminomethane

Reagent2: R2: Aqueous glucose: [Glu]Eta = 100 mg /dL (calibration solution)

Samples: Medicinal solution e.g. 5% glucose serum.

4.Operating mode

1. Using a marker, identify the test tubes from left to right: (Reagent blank, Standard, Sample)

2. Prepare the reagents according to the technical sheet to obtain the Working Reagent , here the reagents are ready to use (prepared).
3. In a beaker and with distilled water, dilute the sample to be analyzed to 1/26. pipette the volumes: (1 ml glucose serum + 25 ml H_2O).
4. Pipette the following volumes into the test tubes:
5.

Reacts \ Tubes	Reactive blank	Calibration	Sample
Reagent 1 (R1) (µl)	2000	2000	2000
Distilled water (µl)	20	-	-
Reagent 2 (R2) (µl)	-	20	-
Sample (µl)	-	-	20

6. Mix and incubate for exactly 5 min at 37°C or 20 min at room temperature (15-25°C).
7. Zero the spectrophotometer using the Reagent Blank.
8. Read and record the Absorbances of the Standard (Abs Calibration) and the sample (Abs sample) before 30 minutes. To calculate the concentration of glucose present in the sample, we apply the following formula:

Example:

$$\begin{cases} [Glu]_{cal} = 100\ mg/dl \longrightarrow Abs_{Cal} = 0{,}392\ abs \\ [Glu]_{Samp} = \mathbf{x} \longrightarrow Abs_{Samp} = 0{,}805\ abs \end{cases}$$

$$[Glu]_{samp} = (Abs_{cal}/\ Abs_{Samp}) \times [Glu]_{Eta} = \mathbf{x}\ \text{mg /dL of glucose in the sample}$$

x = 205 mg /dl = **2,05** g/l this is just an example

5.Questions

1. Calculate the mass concentration after the Dilution of the studied solution [Glu]D in mg/dl.

2. Deduce the mass concentration before dilution (Pure) of the studied solution [Glu]p in mg/dl. Then in g/l

3. Convert this concentration into mass percentage of Glucose found in pure serum.

4. What is your opinion on the quality of this product (5% glucose serum)?

5. If we consider the result calculated in Q1 as a blood glucose level from a patient's fasting blood sample; is he diabetic or not?

Experiment 10
Synthesis of Soap

1.Introduction

Saponification is a process in which a fat molecule is broken down by sodium hydroxide into four smaller molecules; three of the new molecules are soap and one is glycerol. In simple terms saponification is the name for a chemical reaction between an acid and a base to form a carboxylate salt with very long hydrocarbon chains

When you make soap, you mix an oil or fat (which is your acid) with Sodium Hydroxide (which is your base) to form soap (which is a salt). There are many different types of acids that will react with your base and saponify. Your acid could be olive oil, coconut oil or vegetable oil. Each acid has a unique combination of triglycerides (compounds made of three fatty acids attached to a single glycerol molecule) which combines with the base sodium hydroxide differently.

The general reaction, using NaOH, is given by:

$$RCOOR' + NaOH \rightarrow ROH + R'COO^{-}\ Na^{+}$$

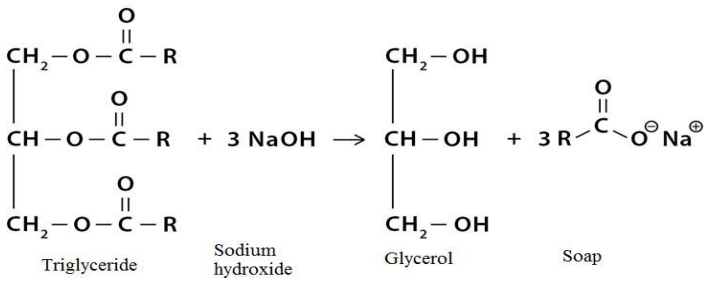


Figure 1: General saponification reaction.

2.Objectives

The aim of this experiment:

1. Describe a procedure for making soap.
2. Explain with the use of equations how soap is formed.
3. Gained an understanding of the saponification process.

3.Experimental

3.1.Materials

Sodium Hydroxide, Vegetable oil (Most oils can be used), distilled water, Sodium Chloride, Beaker, Hot plates / magnetic stirrer, Balance and cricuble.

Health and safety information for materials used Sodium hydroxide: Strong base which must be handled with care (corrosive and caustic). It is important to make sure you are careful when mixing the water and sodium hydroxide. If the chemical comes into contact with your skin, wash the affected area with copious amount of water.

3.2.Procedure

-Weigh 6 g of Sodium Hydroxide; Pour it into a beaker containing 20 mL of distilled water and mix.

-Place a clean beaker on the balance; weigh 5 g of oil. Then to that same beaker added the basic solution. Reweigh, and record the mass.

-Add a small magnetic stir bar to the beaker and heat and stir the mixture on a magnetic stirrer- hotplate. Heat the mixture (with constant stirring) in a bath of boiling water for 45 minutes. While this mixture was heating, dissolve 25g of NaCl into 75 mL of H_2O in an other beaker.

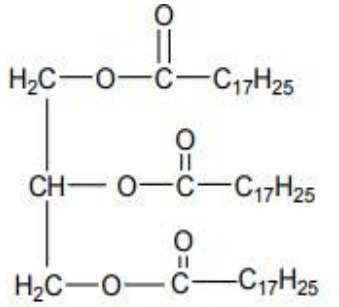
-When the saponification is complete, carefully remove the beaker from the heat. The hot solution was poured into the NaCl solution. This process is called “salting out”. It increases the density of the solution and causes the soap to precipitate and float on the surface of the solution. Lastly, collect the soap curds into a clean, dry watch glass or a small beaker. After drying; weigh the soap.

5.Questions

- Does it smell like any soap that you have used?

-Calculate the percent yield.

-Write the reaction for the saponification of stereol-a triglyceride with sodium hydroxide.

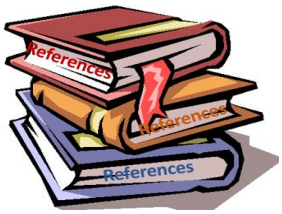


stereol-a triglyceride
(beef fat)

-Why the product of saponification is called a salt?

-Why you added sodium chloride solution at the end of saponification reaction.

-Calculate the molarity, normality and mass concentration of NaOH soltution used in this experiment (Molar Mass = 40g/mol).



[1] Chemistry Lab: From Experiment to Interpretation by Thomas Barilero, Aurélie Deleuze, Matthieu Emond, and Hélène Monin-Soyer. 2013.

[2] Ready-to-Ready Chemistry Lab: Preparatory Classes for PC / BCPST-Véto IUT / BTS / Bachelor's Degree by Thomas Barilero, Aurélie Deleuze, Matthieu Emond, and Hélène Monin-Soyer. 2009.

[3] General Inorganic and Organic Chemistry Lab. 25 Topics, Revised and Corrected Edition by François Souil. 2000.

[4] http://ww3.acpoitiers.fr/sc-phys/tournier/secondes /chimie/verrerie/verrerie.htm

[5] http://www.ostralo.net/materieldelabo/pages/bouchons.htm

[6] Instrumental Techniques for Chemical Analysis. Francis Rouessac, Annick Rouessac. Dunod. 2011.

[7] www.travailler-mieux.gouv.fr: Questions and Answers on CLP / GHS.

[8] www.ineris.fr/ghs-info/: National regulatory assistance service on the CLP regulation.

[9] M. Chavanne, A. Jullien, G.J. Beaudoin, E. Flamand, Experimental Organic Chemistry, 2nd ed., Belin. 1991.

[10] www.inrs.fr and www.9pictos.com: Web file dedicated to the new labeling of chemical products

[11] Mini Manual of General Chemistry - Chemistry of Solutions. Elisabeth Bardez. Dunod. 2008.

[12] 11. Ready-to-use chemistry labs. Thomas Barilero, Aurélie Deleuze, and Matthieu Edmond. Editions Rue d'Ulm. 2009.