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وزارة التعليم العالي والبحث العلمي
Ministry of Higher Education and Scientific Research
جامعة محمد بوضياف – بالمسيلة
Mohamed BOUDIAF University – M'Sila

جامعة محمد بوضياف - المسيلة
Université Mohamed Boudiaf - M'sila

Faculty of Technology

Vice Deanship of Post-Graduation, Scientific Research and External Relations

كلية التكنولوجيا

نيابة العمادة لما بعد التدرج والبحث العلمي والعلاقات الخارجية

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شهادة ادارية

المصادقة على تقارير خبرة للموافقة على مطبوعة بيداغوجية

بعد الإطلاع على تقارير لجنة الخبراء للموافقة على المطبوعة البيداغوجية للأستاذ : **شوق نور الدين – أستاذ محاضر قسم ب**، بالقاعدة المشتركة بكلية التكنولوجيا بجامعة محمد بوضياف بالمسيلة والتي كانت كلها ايجابية ، تمّ تقرير التّالي:

1-المصادقة على تقارير لجنة الخبراء للموافقة المطبوعة البيداغوجية والمعنونة بـ:

Chemistry Practical Work 1
Common Base ST– Cycle License

2- حيث تمّ تشكيل هذه اللجنة بناء على اجتماع المجلس العلمي للكلية المنعقد بتاريخ **2025/11/30** المكونة من السّادة الآتية أسماؤهم:

-**مخلوفي الهاني**، أستاذ محاضر "أ"، جامعة محمد بوضياف- المسيلة

-**سعيد محمد الهادي**، أستاذ محاضر "أ"، جامعة محمد بوضياف- المسيلة

- **لفحل مصطفى**، أستاذ محاضر "أ"، جامعة صالح بوبنيدر - قسنطينة 3

وتمت الموافقة بالاجماع على هذه المطبوعة.

رئيس المجلس العلمي للكلية

المجلس العلمي

د. علي جريوي

PEOPLE’S DEMOCRATIC REPUBLIC OF ALGERIA
MINISTRY OF HIGHER EDUCATION & SCIENTIFIC RESEARCH
UNIVERSITY OF MOHAMED BOUDIAF - M’SILA
FACULTY OF TECHNOLOGY
DEPARTMENT OF COMMON BASE

Chemistry Practical Work 1

First Year License (LMD), Common Base

Dr : CHOUG Noreddine

Academic Year : 2025/2026

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General Introduction

The primary purpose of chemistry laboratory work is to allow students to apply the theoretical concepts acquired in class within an experimental context. These practical sessions play a vital role in developing a concrete understanding of chemical principles and acquiring essential experimental skills.

Laboratory experiments, typically conducted under the supervision of an experienced instructor and in a controlled environment with appropriate safety equipment (such as safety goggles, gloves, and fume hoods), give students the opportunity to handle chemical substances, use precise instruments, measure volumes, perform chemical reactions, conduct analyses, calculate concentrations, carry out titrations, and explore various chemical transformations. Through these activities, students develop practical skills such as accurate handling of equipment, careful observation, data collection, result analysis, problem-solving, and scientific communication.

Laboratory work also reinforces the understanding of fundamental chemistry concepts and offers hands-on experience that complements classroom theory. It helps students acquire transferable skills, such as teamwork, following experimental protocols, implementing proper safety measures, and effectively communicating their findings.

The ultimate objective of laboratory work is to train competent chemists who can conduct experiments, analyze results, and draw meaningful conclusions. These practical abilities are crucial for success in research, the chemical industry, and other chemistry-related fields.

To ensure the smooth running of laboratory sessions, the following general rules should be observed:

- Sessions should be prepared in advance.
- A laboratory report must be submitted at the end of each session, and will be evaluated throughout the semester.
- An unexcused absence will result in a zero for the corresponding report. In the case of an excused absence, no grade will be assigned for that session.

PRACTICAL WORK N^{O}. 01 :

LABORATORY SAFETY

1. Introduction

Laboratory safety is of utmost importance when working with chemicals and potentially hazardous equipment. It is crucial to take the necessary precautions to ensure the health and safety of all individuals in the laboratory. This includes wearing appropriate personal protective equipment such as safety goggles, gloves, and lab coats, as well as handling and storing chemicals correctly. Familiarity with safety data sheets, working in a well-ventilated fume hood, and knowledge of emergency procedures are also essential to minimize the risk of accidents in the laboratory. By adhering to these safety measures, we can create a safe and protected working environment for all researchers and technicians.

2. Objective

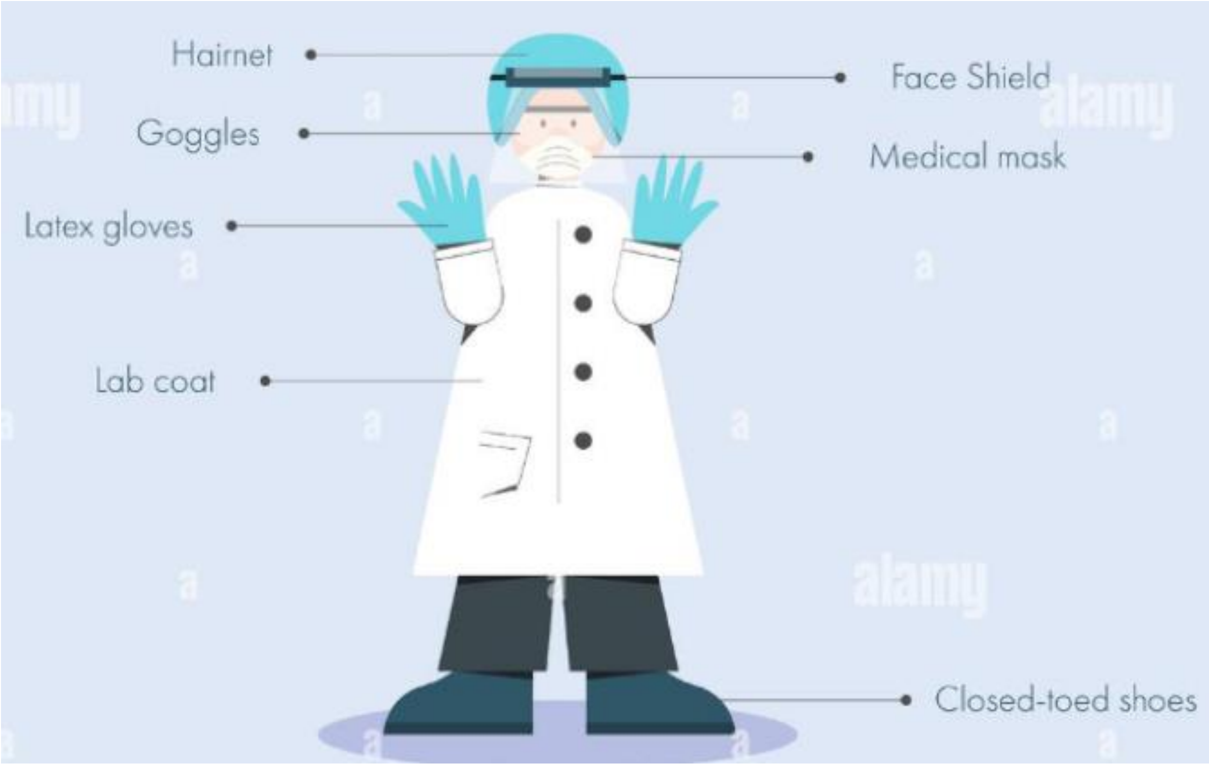
The objective of laboratory safety for students is to guarantee their protection and well-being while working with chemicals and potentially dangerous equipment. The laboratory safety measures aim to minimize the risks of accidents, injuries, and exposure to harmful substances. By comprehending and adhering to the safety rules, students can prevent incidents, minimize potential harm to themselves and others, and promote responsible scientific practices. The ultimate goal is to establish a safe working environment where students can conduct their experiments with confidence and a heightened awareness of safety

3. Laboratory safety

Laboratory safety is of utmost importance when it comes to handling chemicals and laboratory equipment. It is essential to adhere to certain rules to ensure the safety of everyone in the laboratory. Here are some important guidelines to follow in order to maintain laboratory safety

3.1. Personal Protective Equipment (PPE)

In a chemical laboratory, the use of personal protective equipment **(PPE)** is crucial due to the specific risks associated with handling potentially hazardous chemicals. Here are some commonly used **PPE** in a chemical laboratory



Personal Protective Equipment

Lab Coat : A lab coat made of cotton or polyester is typically worn to protect clothing from chemical splashes. It should have chemical-resistant properties and be easy to wash.

Safety Goggles : Safety goggles with lenses resistant to chemicals are essential to protect the eyes from splashes, spills, or chemical vapors.

Respiratory Mask : In laboratories where chemical vapors or particles are present in the air, an appropriate respiratory mask should be worn to protect the respiratory tract.

Protective Gloves: Chemical-resistant gloves are indispensable to shield the hands from direct contact with corrosive or toxic chemicals.

Safety Shoes: Safety shoes with steel toe caps are recommended to safeguard the feet from chemical splashes and the impact of heavy objects.

Protective Apron : To provide additional protection against chemical splashes, a plastic or rubber apron can be worn over the lab coat.

Safety Helmet: In laboratories where there is a risk of heavy objects falling or potential impacts, wearing a safety helmet may be necessary.

Respiratory Protection Equipment : In situations where toxic fumes or gases are present, more advanced respiratory protection devices such as filter respirators or powered air-purifying respirators may be required.

It is crucial to use the appropriate PPE, ensure they are well-fitted, properly maintained, and used in accordance with safety guidelines to effectively protect against potential hazards in the laboratory.

3.2. Collective Protection Equipment (CPE)

In addition to personal protective equipment (PPE), chemical laboratories must also implement collective protection equipment (CPE) to ensure the safety of workers and prevent accidents. Here are some examples of commonly used CPE in a chemical laboratory :

Chemical Fume Hood : A chemical fume hood is a ventilation device that is designed to capture potentially hazardous vapors, gases, or dust that are generated during chemical manipulations. It creates a directed airflow, either towards the outside or through appropriate filtration systems, to protect workers from inhaling harmful substances.

Safety Shower: In the event of contact with corrosive or irritating chemicals, a safety shower allows for quick and effective rinsing of affected body parts. It should be equipped with clean and tepid water to minimize damage caused by these chemicals.

Emergency Eyewash Station: Similar to the safety shower, the emergency eyewash station provides a rapid means of rinsing the eyes in case of contact with chemicals or other irritating substances. It is crucial that the eyewash station is easily accessible and functions properly at all times.

Containment Systems: Chemical laboratories may be equipped with containment systems such as glove boxes or glove bags to safely handle potentially hazardous chemicals. These systems help reduce the risk of exposure by confining the substances within a secure area.

Gas Detection Systems: Gas detectors can be installed to continuously monitor the presence of toxic or flammable gases in the laboratory air. In the event of detecting a dangerous gas concentration, alarms can be triggered to alert workers and initiate evacuation measures.

Signage and Labeling: It is important to use appropriate signage, safety labels, and floor markings to identify hazardous areas, safety devices, and emergency procedures in the laboratory.

Training and Emergency Procedures : In addition to collective protection equipment, proper training on the use of this equipment and emergency procedures to follow in the event of a chemical accident is crucial to ensure worker safety.

The effective utilization of these collective protection equipment in conjunction with ***PPE*** can significantly contribute to reducing the risks of chemical accidents and safeguarding the health of workers in a chemical laboratory.

3.3. Some General Rules for Laboratory Safety

- Do not eat or drink in the **Laboratory**.
- Do not run inside the **Laboratory**.
- Do not throw anything in the **Laboratory**.
- Use chemicals while standing.
- Always keep walkways, exits, and safety equipment clear.
- Keep flammable materials away from fire or heat.
- Label every chemical container.
- Close all containers tightly after use.
- When pouring chemicals, pour away from the label.
- Do not return unused chemicals to the original container unless your teacher says so.
- It is better not to wear jewelry in the **Laboratory**.
- Do not sit on **Laboratory** tables.

4. A Chemical Product Label: Importance and Key Elements

A chemical product label is a vital source of information that provides essential details about a specific chemical substance. It is usually affixed to the container such as a bottle, drum, or canister—and serves a critical role in ensuring safe handling, storage, and usage. Proper interpretation of the label is fundamental for preventing accidents, protecting health, and complying with legal requirements.

4.1 Reading a Chemical Product Label

Being able to read and interpret a chemical product label is an essential safety skill. The label contains standardized elements designed to convey hazards and precautions clearly. Below are the key components you will typically find:

1. **Warning Statement (Signal Word)**

- **Definition:** A standardized term that indicates the **severity level** of a hazard on a chemical product label or Safety Data Sheet (SDS).
- **Types:**
  - **DANGER:** Used for the **most severe hazards**, which can cause death or irreversible harm.
  - **WARNING:** Indicates **less severe hazards**, but still capable of causing injury or material damage.
- **Example:**
  "DANGER: Toxic if inhaled."

Why it matters: The signal word immediately alerts the user to the seriousness of the risk and prompts appropriate caution.

2. **Product Identifier**

- **Purpose:** Ensures **clear and unambiguous identification** of the chemical substance.
- **Common Elements:**
  - Official **chemical name** (e.g., *Sulfuric Acid*).
  - **CAS number** (Chemical Abstracts Service) – a unique numeric identifier (e.g., *7664-93-9*).
  - **Trade name**, commercial designation, or **batch/lot number**.
- **Importance:** Critical for **inventory control**, **traceability**, and **emergency response**.

3. **Supplier / Manufacturer Identification**

The label and SDS must include **contact details** for the responsible party:

- Full **name of manufacturer or importer**.
- Complete **address**, including country.
- **Telephone number** with international dialing code if necessary.
- **Emergency contact number** (ideally available 24/7).

Why it matters: This ensures that users or emergency responders can **quickly obtain technical support** in case of spills, exposure, or accidents.

4. **Precautionary Measures**

- **Definition:** Instructions aimed at **reducing or eliminating risk** during product handling.
- **Examples:**
  - **Storage:** "Store in a cool, well-ventilated area away from heat sources."
  - **Handling:** "Wear chemical-resistant gloves (e.g., nitrile) and protective eyewear."
  - **First Aid:** "If skin contact occurs, rinse immediately with plenty of water."

Objective: Minimize **human health risks**, prevent **environmental contamination**, and ensure **safe operations**.

5. **Hazard and Precautionary Statements**

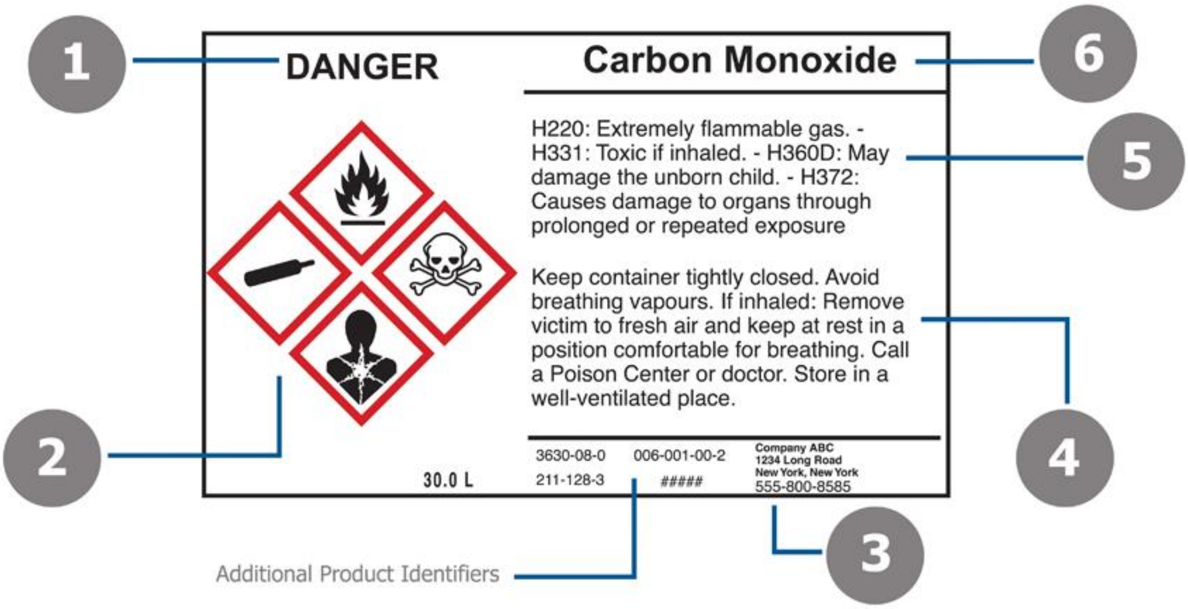
- **Hazard Statements (H-statements):** Standardized phrases describing the nature and degree of hazard (physical, health, or environmental).
  - Examples:
    - **H225:** Highly flammable liquid and vapor.
    - **H314:** Causes severe skin burns and eye damage.
- **Precautionary Statements (P-statements):** Recommendations on how to **reduce or control exposure** and manage risks.
  - Examples:
    - **P210:** Keep away from heat, hot surfaces, sparks, open flames.
    - **P305+P351+P338 :** IF IN EYES : Rinse cautiously with water for several minutes.

Purpose: These statements are **mandatory under the GHS (Globally Harmonized System)** to standardize safety communication worldwide.

6. **Hazard Pictograms (GHS)**

- **Definition:** Standardized graphic symbols framed in a red diamond shape, representing the **type of hazard**.
- **Examples of pictograms and meanings:**
  - **Flame:** Flammable substances.
  - **Skull and Crossbones:** Acute toxicity (fatal or toxic).
  - **Corrosion:** Causes skin corrosion/burns or eye damage.

- **Health Hazard:** Carcinogenicity, mutagenicity, respiratory sensitization.
- **Environment:** Hazardous to aquatic life.



4.2. The topic of hazard symbols and their meanings

Hazard symbols, also known as warning symbols or hazard pictograms, are graphical representations used to communicate specific hazards associated with chemicals or substances. These symbols are standardized internationally to ensure consistency in safety communication across different regions and industries.

Corrosive Hazard Symbol

This symbol warns of substances capables of causing severe damage to skin, eyes, or tissues upon contact. It typically depicts an image of a hand or acid drop on the skin.

Toxic Hazard Symbol

This symbol indicates substances that can cause toxic effects, ranging from mild irritations to severe damage or even death upon exposure. It is commonly represented by a skull and crossbones.

Flammable Hazard Symbol

This symbol warns of substances that can easily ignite and burn rapidly. It is represented by a flame, indicating the potential fire hazard.

Explosive Hazard Symbol

This symbol signifies substances that are prone to causing explosions under certain conditions such as heat, shock, or friction. It is depicted by an explosion, indicating the risk of explosion.

Health Hazard Symbol

This symbol warns of substances that can pose risks to health, such as irritations, allergies, respiratory issues, or carcinogenic effects. It is typically represented by an exclamation mark, highlighting the general health hazard

Oxidizing Hazard Symbol

This symbol indicates substances that can cause or promote the combustion of other materials. It is represented by a circle with flames inside, signifying the potential for fire.

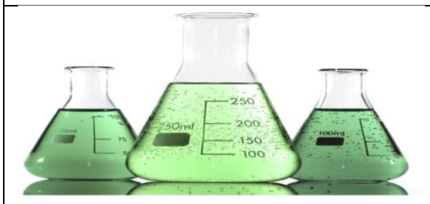
5. Chemistry Safety & First Aid Guide

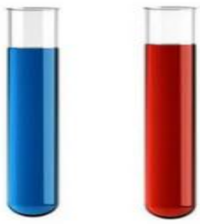
Danger	Safety rule	First aid actions
Ingestion of a chemical	Never eat, drink, or taste chemicals. Always handle chemicals properly and use appropriate PPE.	If someone ingests a chemical, call poison control or emergency services. Do not induce vomiting unless directed by professionals.
Eye exposure	Always wear appropriate safety goggles when handling chemicals or objects that may cause splashes.	Rinse thoroughly with clean water for at least 15 minutes. Seek immediate medical attention.

Thermal burn	Avoid heat sources and handle hot objects with caution.	Cool the affected area with cold water for at least 10–20 minutes. Cover with a clean dressing. Seek medical attention if necessary.
Chemical burn	Use appropriate PPE when handling corrosive chemicals.	Rinse affected area with water for at least 20 minutes. Remove contaminated clothing unless stuck to skin. Seek immediate medical attention.
Cut	Use sharp tools with caution and avoid sudden movements.	Clean the wound with water and soap. Apply pressure and cover with a clean dressing. Seek medical attention if necessary.
Fire	Know emergency exits, evacuation procedures, and fire extinguisher locations.	Alert others, activate fire alarm, evacuate. Use fire extinguisher only if trained and safe to do so.
Inhalation of an irritating or toxic gas	Work in a well-ventilated environment and use respiratory protection when needed.	Move to fresh air immediately. Call emergency services if necessary. Seek medical attention if symptoms persist.

6. Glassware used in laboratory chemistry

the glassware used in chemistry, an essential component in laboratory work, Glassware is used for precise measurements, mixing, reactions, and analysis in the field of chemistry. Here are some common types of glassware used in chemistry:

	**Beaker :**Beakers are basic containers used to hold samples and reagents. They come in various sizes and have a wide opening, making it easy to add materials and mix liquids
	**Erlenmeyer :**Erlenmeyer flasks have a narrow neck and a flat bottom. They are used for swirling, storing, and heating liquids. These flasks are commonly used when a container needs to be sealed, as they can easily accommodate stoppers or be covered with parafilm

	**Volumetric flasks :** is a type of laboratory glassware used for precise measurement and preparation of solutions . It is designed to contain a specific volume of liquid at a particular temperature. The flask has a flat bottom and a long neck with graduation marks indicating the volume contained in the flask .
	**Graduated Cylinder :** A graduated cylinder is a common laboratory glassware used to measure the volume of liquids accurately. It consists of a narrow cylindrical container with calibrated markings along its side.
	**Graduated pipettes :**Graduated pipettes are another type of laboratory glassware used for accurately measuring and transferring liquids. They are typically made of glass or plastic and have a tapered tube with a graduated scale marked along its length. Graduated pipettes are commonly used in biology, chemistry, and medical laboratories for tasks such as preparing solutions, transferring precise volumes of reagents, and performing serial dilutions. They offer a high level of accuracy and reproducibility in liquid handling procedures
	**Funnel:**A funnel is a laboratory tool with a wide mouth and a narrow stem. It is used to safely transfer liquids or fine-grained substances into containers with small openings, preventing spills. In combination with filter paper, it is also used for filtration to separate solids from liquids.
	**Test tube :** test tube is a cylindrical, transparent, or translucent vessel made of glass or plastic, typically open at one end and rounded at the other. They come in various sizes and are used in laboratory settings for holding, mixing, or heating small quantities of liquid or solid substances. Test tubes are versatile tools in scientific experiments and are commonly found in chemistry, biology, and medical laboratories.

	**Burette** is a long, narrow, graduated glass tube with a stopcock at the bottom, used for accurately measuring and dispensing precise volumes of liquid in laboratory settings. Burettes are commonly used in titration experiments, where a solution of known concentration (the titrant) is added gradually to another solution of unknown concentration (the analyte) until a chemical reaction between the two is complete.

7. Evaluation

1. Why should one wear a lab coat in a chemistry laboratory?

2. What are the different pieces of information that can be found on the label of a chemical product?

3. Provide the meaning of the following pictograms?

4- What should you do in the following circumstances? Eye exposure Answer:

5- Identify the laboratory glassware to follow :

8. Conclusion

It is essential that students understand the basic safety rules to avoid accidents and injuries in the chemical laboratory.

Students must learn to handle chemicals safely to minimize the risk of accidents and injuries.

Adequate understanding of chemical labels allows students to recognize the potential hazards associated with each product and take appropriate precautions.

Proper use of personal an

Students must be prepared to react quickly and effectively in case of an emergency in the chemical laboratory to minimize the risks to their safety and that of others.

PRACTICAL WORK N^{O}. 02 :

PREPARATION OF SOLUTIONS

1. INTRODUCTION

Experiment 2 provides essential information about solutions in chemistry. It begins by explaining the concept of a solution—a homogeneous mixture of two or more pure substances where the substances do not chemically interact with each other. This definition emphasizes the uniform distribution of components within the mixture. Additionally, it clarifies the roles of solute and solvent. The solute is the substance being dissolved, while the solvent is the substance in which the solute dissolves. Water is highlighted as a common solvent due to its ability to dissolve a wide range of substances.

2. OBJECTIVE

- To recognize and correctly use laboratory equipment and tools for preparing chemical solutions.
- To learn how to prepare a solution from a solid compound
- To learn how to dilute an acid solution in order to obtain the desired concentration.

3. SOME DEFINITIONS

3.1. Solution

A solution is a homogeneous mixture composed of two or more substances. It is typically made up of a solute dissolved in a solvent. Solutions are widely used in various fields of chemistry and are commonly encountered in everyday life. For example, saltwater is a solution in which salt (solute) is dissolved in water (solvent).

3.2. Solute

The solute is the substance that is dissolved in a solution. It is generally present in lesser quantities compared to the solvent. In a saltwater solution, for instance, the salt is the solute as it is dissolved in the water.

3.3. Solvent

The solvent is the substance that dissolves the solute in a solution. It is the component of the solution present in greater quantity. In our example of saltwater solution, the water is the solvent as it dissolves the salt

4. DIFFERENT EXPRESSIONS OF CONCENTRATION

4.1. Molar concentration

This is the number of moles of solute dissolved in one liter of solution. It is expressed in moles per liter (mol/L) or as molarity (M). The formula is:

$$C = \frac{n_{\text{solute}}}{V_{\text{solution}}}$$

Where C is the molar concentration, n_{solute} is the number of moles of solute, and V_{solution} is the volume of the solution.

4.2. Mass concentration

This is the mass of solute dissolved in a unit volume of solution. It is usually expressed in grams per liter (g/L) or as a percentage (% by weight). The formula is:

$$C = \frac{m_{\text{solute}}}{V_{\text{solution}}}$$

Where C is the mass concentration, m_{solute} is the mass of the solute, and V_{solution} is the volume of the solution.

4.3. Normality

Normality is a measure of concentration used in chemistry, specifically in aqueous solution, and is defined as the number of equivalents of solute present in one liter of solution. An equivalent of solute is the amount of solute that can furnish or accept one mole of hydrogen ions (H^+) in a chemical reaction.

The general formula to calculate the normality (N) of a solution is as follows:

$$N = \frac{n}{V}$$

Where:

N : is the normality of the solution (in equivalents per liter, eq/L or N)

n : is the number of equivalents of solute

V : is the volume of the solution (in liters)

For some strong acids and bases, normality is the same as molarity (M), but for weak acids and bases or solutions containing multiple ionizable groups, it may differ.

It is important to note that normality can differ from molarity for acids and bases that can donate or accept more than one H^+ or OH^- ion per mole in solution. For example, for sulfuric acid (H_2SO_4), which can furnish two H^+ ions per mole, a 1 M solution would have a normality of 2 N.

When preparing solutions from solutions of different normalities, you can use the following equation to balance the normalities:

$N1 * V1 = N2 * V2$

Where N1 and N2 are the normalities of the stock solution and the diluted solution, respectively, and V1 and V2 are the volumes of these solutions. This allows you to calculate the volume of stock solution needed to prepare the diluted solution at the desired concentration.

5. RELATION BETWEEN NORMALITY AND MOLARITY

The relationship between molarity (M) and normality (N) provides two ways to measure the concentration of a solution. You already have a good foundation with the formula

N = M * Z, But we can elaborate further.

Molarity (M) : is the number of moles of solute per liter of solution. It's a simple and direct measure of the amount of substance present.

For example, a 0.05 M NaCl solution contains 0.05 moles of NaCl per liter.

Normality (N) : on the other hand, is a measure of concentration that takes into account the reactive capacity of the solute. It expresses the number of gram equivalents of solute per liter of solution. A gram equivalent is a measure of the amount of a substance that can react with a specific number of other substances. This is where the Z factor comes in.

The Z factor (or equivalence factor) represents the number of reactive units per molecule of solute. This number depends on the chemical reaction being considered. Here are some examples to clarify:

ACIDES : Z corresponds to the number of H^+ protons that the acid can release during a neutralization reaction.

- For HCl (hydrochloric acid), Z = 1 because it releases a single proton.
- For H_2SO_4 (sulfuric acid), Z = 2 because it can release two protons.

BASES : Z corresponds to the number of OH^- hydroxide ions that the base can release.

- For NaOH (sodium hydroxide), Z = 1.
- For $Ba(OH)_2$ (barium hydroxide), Z = 2.

REDOX REACTIONS : Z corresponds to the number of electrons transferred per molecule of solute. In the example you provided,

$[MnO_4^- + 5e^- + 8OH^- \rightarrow Mn^{2+} + 4H_2O]$, Z = 5.

because 5 electrons are transferred.The apparent standard potential, defined by the Nernst equation, is influenced by pH and the activities of the oxidizing and reducing species.

SALTS : For salts, Z is more complex and depends on the charge of the metal ions. In the case

Of $Al_2(SO_4)_3$, Z = 6 because there are two Al^{3+} ions (total charge +6). The solubility of salts, like AgCl, is related to its solubility product Ks, which is related to the solubility s.The concentration of ions in solution depends on the stoichiometric coefficients.

In summary, normality is a more specific measure than molarity because it considers the reactivity of the solute in a given reaction. The formula **N = M * Z** allows for the conversion of molarity to normality, but it is crucial to correctly determine the value of Z based on the reaction context. Temperature and pressure can also influence the solubility of gases and the concentration of species in solution.

6. EXPERIMENTAL PART

OBJECTIVE OF THE EXPERIMENT

1. Recognising and using the equipment and tools for preparing solutions.

2. How to prepare a solution from sodium hydroxide (*NaOH)*∗ by dissolving.

3. How to dilute Hydrochloric acid (HCI) solution.

6.1. Preparation of solution from solid

Materials - Products Used

Materials	
Watch glass	Pipette bulb/filler (propipette)
Spatula	Beaker
Balance	Volumetric flasks (100 mL and 250 mL)
Funnel	Wash bottle (distilled water)

***Warning* :**

Remember to consider any safety precautions necessary while handling solid solutes and working in the laboratory. Wear appropriate personal protective equipment (PPE) such as gloves, a lab coat, and safety goggles.

I. PREPARATION OF A SOLUTION BY DISSOLVING A SOLID COMPOUND

We want to prepare 100 mL of an aqueous magnesium sulfate solution with a concentration C of 0.1 mol. L^{-1}.

1. Determine the molar mass of magnesium sulfate :

Magnesium sulfate's chemical formula is $MgSO_4$.

Mg: 24.31 g/mol S: 32.07 g/mol O: 16.00 g/mol .

2. Calculate the required mass (m) of the product: We use the formula :

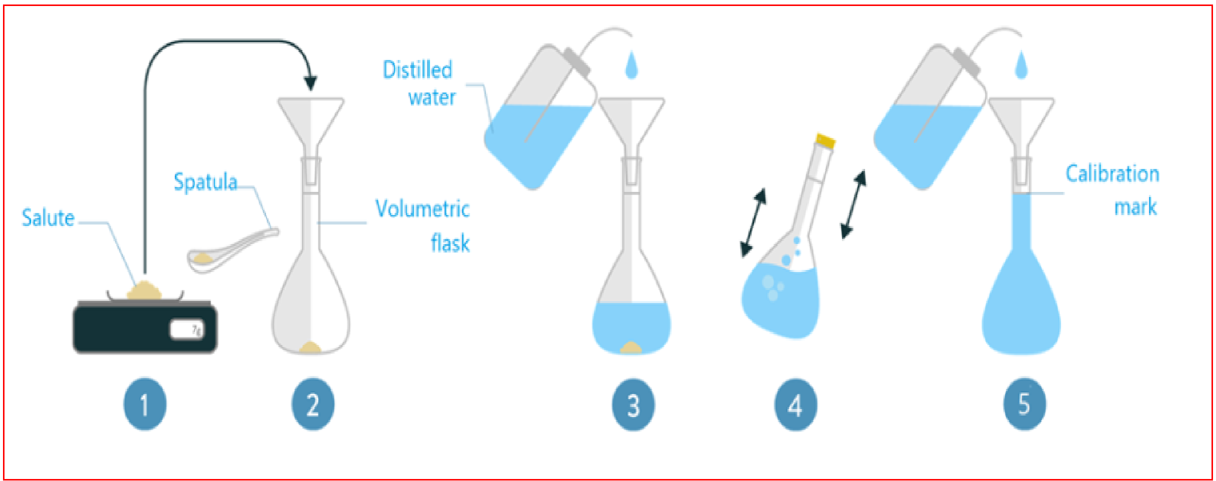
   Moles = Concentration (C) * Volume (V)

3. **Preparation Steps:**

- Weigh the calculated mass (1.2038 g) of solid magnesium sulfate using a clean, dry spatula and a watch glass.
- Using a funnel, transfer the solid into a 100 mL volumetric flask. Rinse the watch glass and funnel with distilled water, collecting the rinsings in the flask.
- Fill the volumetric flask about three-quarters full with distilled water. Stopper the flask and swirl until the $MgSO_4$ crystals are completely dissolved.
- Carefully add distilled water until the bottom of the meniscus aligns with the graduation mark on the flask (hold the flask at eye level to avoid parallax error).
- Stopper the flask securely and invert it several times to ensure thorough mixing. You have now prepared the "stock" solution of $MgSO_4$.

4. **Determine the uncertainty in the molar concentration :**

Uncertainty comes from various sources: the balance's precision, the volumetric flask's tolerance, and the purity of the $MgSO_4$. A simplified approach, assuming the weighing error is dominant, is described in the previous response.



How to prepare solutions from solid

II. PREPARATION OF A SOLUTION BY DILUTION

We want to prepare an aqueous solution (F) of magnesium sulfate with a molar concentration C_f= 0.02 mol. L^{-1}. We only have the previously prepared stock solution with a concentration

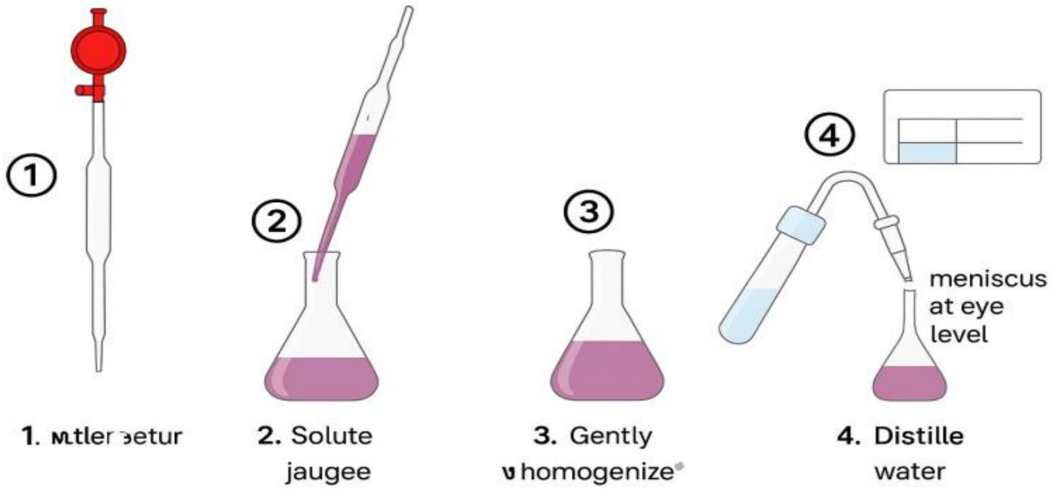
C = 0.1 mol. L^{-1} and a 250 mL volumetric flask.

1. Determine the volume (V) to take from the stock solution We use the dilution formula

$$C_1V_1 = C_2V_2$$

2. Experimental protocol for preparing solution F.

- Using a pipette, accurately measure 50 mL of the stock $MgSO_4$ solution.
- Transfer the measured volume into the 250 mL volumetric flask.
- Add distilled water to the flask, filling it about three-quarters full.
- Swirl the flask to mix the solution.
- Carefully add distilled water until the bottom of the meniscus aligns with the graduation mark on the flask.
- Stopper the flask securely and invert it several times to ensure homogeneity.



How to prepare solutions through dilution

III. PREPARATION OF A SOLUTION FROM A COMMERCIAL SOLUTION

We want to prepare 250 mL of a 0.1 mol/L hydrochloric acid (HCl) solution from a commercial HCl solution with a density of 1.18 g/mL and 36% concentration (by mass).

1. Calculate the molar concentration of the commercial solution.

2. Calculate the volume (V) of commercial HCl to be taken: Using $C_1V_1 = C_2V_2$.
3. Preparation steps (as described) : These steps emphasize the important safety precaution of adding acid to water, along with the use of a cooling bath to manage the exothermic reaction.

PRACTICAL WORK N^{O}. 03 :

ACID-BASE TITRATION

1. INTRODUCTION

Acid-base titration is a quantitative analysis technique in chemistry used to determine the unknown concentration of an acid or base solution. This method relies on a controlled chemical reaction between the solution being analyzed and another solution of a precisely known concentration.

2. OBJECTIVE

- Understand the principles of acid-base titration, including the concept of the equivalence point and detection methods (colorimetric and PH-metric).
- Put these principles into practice to determine the concentration of a strong acid solution (HCl) by titration with a strong base (NaOH) of known concentration.
- Apply the titration technique to determine the acidity of commercial vinegar (a weak acid solution, CH3COOH) by titration with a strong base (NaOH) and compare the experimental result with the value indicated on the label.

3. DEFINITIONS OF AN ACID AND A BASE

Understanding acids and bases is fundamental to titration. The most commonly used theory in this context is the Brønsted-Lowry theory.

- **Brønsted-Lowry Acid** : An acid is a chemical species (a molecule or ion) capable of donating one or more protons (H^+ ions).
- **Brønsted-Lowry Base** : A base is a chemical species capable of accepting one or more protons (H^+ ions).

When an acid donates a proton, it forms its conjugate base. Conversely, when a base accepts a proton, it forms its **conjugate acid**. These two species form what is known as an **acid/base** pair, denoted as Acid/Base.

Example :
In water, hydrochloric acid (HCl) donates a proton to water (H_2O).
$HCl + H_2O \rightarrow Cl^- + H_3O^+$

- HCl is the acid (proton donor).

- H_2O is the base (proton acceptor).
- Cl^- is the conjugate base of HCl (forming the HCl/Cl^- pair).
- H_3O^+ (hydronium ion) is the conjugate acid of H_2O (forming the H_3O^+/H_2O pair).

4. PRINCIPLE OF TITRATING AN ACID WITH A BASE

The goal of a titration is to determine the concentration of a solution (called the **analyte** or **titrand**) by reacting it with another solution of known concentration (called the **titrant**). The core reaction in an acid-base titration is a **neutralization reaction**. During this reaction, an acid and a base react to form a salt and water.

General equation :

Acid + Base → Salt + Water

To be suitable for a titration, this reaction must be:

- **Complete** : It must consume virtually all of the reactants.
- **Fast**:The equilibrium must be reached almost instantaneously after each addition of the titrant.
- **Unique** : There should be no interfering side reactions.

The key moment in the titration is the **equivalence point**. This is the precise point where the reactants (the acid and the base) have been mixed in the stoichiometric proportions of the reaction. In other words, the amount of titrant added is just enough to completely react with the initial amount of the analyte.

At the equivalence point, for a reaction where one mole of acid reacts with one mole of base, the following relationship holds true:

$$C_A * V_A = C_B * V_E$$

Where:

- C_A is the concentration of the acid (often the unknown).
- V_A is the volume of the acid sample.
- C_B is the known concentration of the base.
- V_E is the volume of base added to reach the equivalence point (the volume we aim to measure).

5. ACID-BASE TITRATION METHODS

To determine the equivalence volume (V_E), one must monitor the progress of the reaction. There are two main methods for this.

5-1- Colorimetric Titration (with a Color Indicator)

This method is quick and does not require complex equipment.

- **Principle** : A few drops of a **color indicator** are added to the analyte solution. A color indicator is a weak acid/base pair where the acid form and the base form have different colors.
- **Detection**: The color of the indicator depends on the PH of the solution. When the pH changes sharply around the equivalence point, the indicator changes color. This color change signals the **end point** of the titration.
- **Choosing the Indicator** : For the measurement to be accurate, one must choose an indicator whose color change range (known as the transition range) includes the expected PH at the equivalence point. For example, bromothymol blue (BTB) is suitable for a strong acid/strong base titration where the PH at the equivalence point is 7.

  Examples of colorimetric indicators:

Indicator	Acid form color	PH transition range	Basic form color
Methyl Orange	Red	3.1 - 4.4 (Orange)	Yellow
Bromothymol Blue	Yellow	6.0 - 7.6 (Green)	Blue
Phenolphthalein	Colorless	8.2 - 10.0 (Pale Pink)	Pink

5-2- PH-metric Titration (Potentiometric Monitoring)

This method is more precise and provides more information.

- **Principle**: A PH meter electrode is immersed in the analyte solution to continuously measure the PH throughout the addition of the titrant.
- **The Titration Curve**: A graph of PH versus the volume of titrant added (PH = f(V)) is then plotted. This curve typically shows a very slow change in pH at the beginning, followed by a very sharp increase (or decrease) called the **PH jump**, and then another stabilization.
- **Detecting the Equivalence Point** : The equivalence point is located in the middle of this PH jump. It can be determined graphically with high precision using several methods:

- **The Parallel Tangents Method**: Two tangents to the curve are drawn, parallel to each other, on either side of the pH jump. A third line is then drawn equidistant from the first two. The intersection of this third line with the titration curve gives the equivalence point.
- **The Derivative Method**: The derivative of the pH with respect to the volume (dPH/dV) is calculated and plotted. This derivative curve shows a peak (an extremum) whose x-coordinate corresponds exactly to the equivalence volume, V_E. This is the most accurate method.

6. EXPERIMENTAL PART :

6-1- Titration of a Strong Acid with a Strong Base

Titration of hydrochloric acid HCl of unknown concentration with sodium hydroxide NaOH of known concentration.

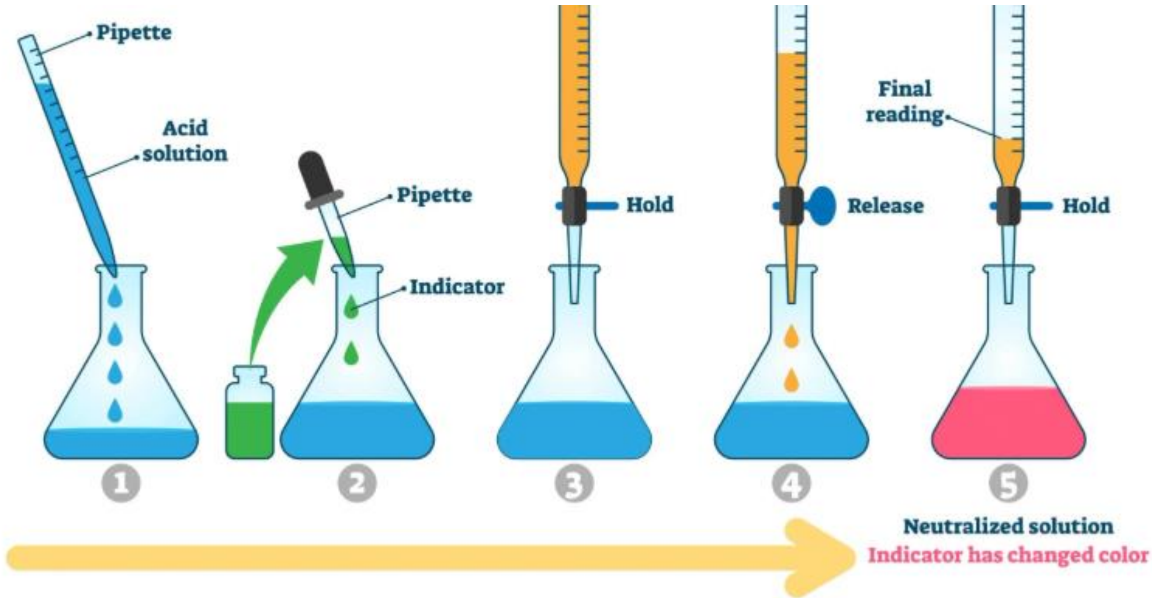
1. Materials and Products:

Materials	
Graduated burette 25ml	Bottle of bromothymol blue
Beaker	PH meter
Graduated cylinder	NaOH solution (0.1 mol/L),
Wash bottle with distilled water	Diluted HCl solution
Magnetic stirrer	Volumetric pipette

2. Procedure:

- Fill the burette with NaOH (0.1 mol/L); remember to prime it and set the zero mark correctly.
- Using a volumetric pipette, accurately measure 10 mL of the diluted HCl solution of unknown concentration (this solution is diluted 10 times). Pour it into a beaker with a few drops of the bromothymol blue indicator.
- Using the burette and a PH meter, add the NaOH solution drop by drop, recording the PH of the solution every 2 mL until a color change occurs (record the volume V_E of NaOH added). Fill in the following table

V_B	0	2	4	6	8	10	12	14	16	18
PH										



Evaluation :

- What is the role of the colored indicator?
- Why was the bromothymol blue indicator chosen?
- Give the volume of sodium hydroxide added at the equivalence point : V_E.
- Give the chemical reactions with all the entities present in the medium: before, after, and at the equivalence point.
- Calculate the concentration C_A and what is the PH obtained at the equivalence point?

- Plot the graph PH = f(V) on graph paper.
- Using the parallel tangent method, determine the coordinates of the equivalence point.
- With the coordinates of the equivalence point, deduce the concentration of the initial acid solution, calculate its pH, and does it agree with the value you measured?
- What does a transition zone mean to you?

6-2- Titration of a Weak Acid with a Strong Base

White vinegar is essentially an aqueous solution of ethanoic (or acetic) acid. Commercial concentrations are expressed in degrees of acidity. The degree of acidity (D°) of white vinegar is expressed by the mass, in grams, of pure ethanoic acid contained in 100g of vinegar. Each bottle indicates the degree of acidity of the commercial vinegar. 1% acidity corresponds to 1g of ethanoic acid in 100g of vinegar solution. Thus, the percentage of acidity corresponds to a mass percentage.

For a volume of 100g of vinegar (CH3-COOH) (Ac: acetic acid):

$m_{AC} = n_{AC} * M_{AC} = C_A * V_A * M_{AC}$ (according to the formula $n = C * V$)

Where:

m_{AC} : is the mass of acetic acid

n_{AC} :is the number of moles of acetic acid

M_{AC} :is the molar mass of acetic acid

C_A :is the concentration of acetic acid

V_A : is the volume of the vinegar solution

We also have for vinegar: $\boldsymbol{\rho} = \frac{\boldsymbol{m}}{\boldsymbol{V}}$ (density = mass/volume)

Therefore:

$$m_{AC} = \frac{m}{\rho} * C * M_{AC} = \frac{100}{\rho} * C * M_{AC} \ (g)$$

Degree of acidity:

$$D^0 = m_{AC} = \frac{100}{\rho} * C * M_{AC}$$

M (CH3-COOH) = 60 g.mol-1 , ρ = 1,01 kg.L-1

Example : A vinegar at 9° contains 9g of pure acetic acid (ethanoic acid) per 100g of vinegar solution.

1) Objective of the Experiment:

Determine the degree of acidity of a commercial white vinegar by titration.

2. Materials and Products:

Materials	
Burette	Graduated cylinder
Beaker	Magnetic stirrer
Wash bottle with distilled water	NaOH solution (0.1 mol/L),
Wash bottle with distilled water	Commercial vinegar (CH3COOH)
Erlenmeyer flask	Bottle of phenolphthalein.

3) Procedure:

- Dilute the vinegar (10-fold) by preparing 50mL of diluted solution.
- Using a pipette, take 10mL of solution A (diluted vinegar) into an Erlenmeyer flask.
- Add 3 drops of phenolphthalein.
- Fill the burette with NaOH solution of molar concentration C_B = 0.1 mol/L.
- Place the Erlenmeyer flask under the burette. Swirl to homogenize the mixture.
- Add NaOH dropwise from the burette until a persistent pink color appears.
- Record the volume V_B of sodium hydroxide corresponding to the color change zone.
- Repeat the experiment 2 times.

4) Evaluation :

- Write the balanced chemical equation for the reaction:
- Record the volume V_B of NaOH added and determine the concentration C_A of CH3COOH.
- Calculate the initial molar concentration C_0 of the commercial ethanoic acid.
- Calculate the degree of acidity (D) of the vinegar and compare this result to the one on the label of the commercial vinegar bottle.
- Plot the graph representing the PH = f(V_B) on graph paper (using the same table as before).
- Determine the volume V_E and the PH at the equivalence point using the graph.
- From the colored indicators listed in the previous table, which one is best suited to detect the equivalence point ?

PRACTICAL WORK N^{O}. 04 :

ACID-BASE TITRATION – CONTROL OF MILK FRESHNESS AND CALCULATION OF THE DORNIC DEGREE.

1- INTRODUCTION

The quality of cheese is directly linked to the quality of the milk used in its production. Since milk serves as the primary raw material, any variation in its composition or freshness has a significant impact on the texture, flavor, and safety of the final dairy product.

Milk is a highly perishable biological fluid, rich in nutrients such as proteins, fats, and carbohydrates, which makes it an excellent medium for microbial growth. Therefore, evaluating its freshness before processing is essential. One of the most widely used indicators of milk freshness is its acidity, which is measured in Dornic degrees (°D).

By definition, 1 °D corresponds to 0.1 g of lactic acid per liter of milk. Lactic acid is not initially present in fresh milk but results from the fermentation of lactose by naturally occurring bacteria.

- Fresh milk generally presents an acidity between 15 and 18 °D and contains approximately 5% lactose.
- As time passes, microorganisms in milk convert lactose into lactic acid, leading to an increase in acidity and a decrease in pH.
- When acidity exceeds 37 °D, casein proteins in milk undergo coagulation, resulting in curd formation. This phenomenon indicates spoilage and unsuitability for cheese production.
- For milk to be considered fresh and safe for processing, its acidity must not exceed 18 °D.

2-FORMULA FOR DORNIC DEGREE

$$D = \frac{C * V_{eq} * M_{AC}}{0.1 * V_0}$$

Where:

C : concentration of NaOH solution

V_{eq} : volume of NaOH at equivalence

M_{AC} : molar mass of lactic acid

V_0 : volume of milk used

3-OBJECTIVES

- To determine the freshness of milk by calculating its lactic acid concentration, then express it in Dornic degrees (°D).

- To determine the lactic acid concentration of processed milk products (fresh cheese or natural yogurt).

4- EXPERIMENTAL PART :

- **Materials and Products :**

Materials	
Burette	Volumetric pipette (25 mL or 10 mL)
Pipette filler	Distilled water
Phénolphtaléine (1% solution in 95% ethanol)	PH meter
Bromothymol blue	fresh milk
Graduated cylinder	Magnetic stirrer with bar
2 Beakers “milk” and “cheese	NaOH solution (0.05 and 0.25 mol/L)
Beaker labeled “waste”	fresh cheese or natural yogurt

4-1-Part 1 — TITRATION OF LACTIC ACID IN MILK

1. Fill the burette with 0.05 mol/L NaOH solution.
2. Adjust liquid level to zero by draining excess into the “waste” beaker.
3. Pipette 20 mL milk into the “milk” beaker.
4. Add 10 drops of phenolphthalein.
5. Insert the magnetic bar.
6. Place the beaker on the stirrer under the burette.
7. Adjust stirring speed to avoid splashing.
8. Insert the pH meter into the milk.
9. Add NaOH gradually (1 mL at a time), record PH after each addition.
10. Identify the equivalence point : persistent color change + sharp rise in PH.

4-2-PART 2 — TITRATION OF LACTIC ACID IN CHEESE/YOGURT

1. Fill the burette with 0.25 mol/L NaOH solution.
2. Adjust liquid level to zero by draining excess into the “waste” beaker.
3. Measure 20 mL of cheese/yogurt.
4. Homogenize in the “cheese” beaker. Dilute to 75 mL with distilled water if too thick.
5. Add 10 drops of bromothymol blue.
6. Insert the magnetic bar and place on the stirrer.
7. Insert the PH meter.
8. Add NaOH gradually; drop by drop near equivalence.
9. Stop when a persistent color change appears with a sharp PH increase.
10. Record the equivalence volume (Veq).

- **OBSERVATIONS**

To complete during experiment)

V_b (mL)	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
PH																

- **INSTRUCTIONS** :

- Plot the titration curve: **PH = f(V_b)**.
- Determine Veq.
- Compare equivalence with indicator color change.

- **ANALYSIS**

- Draw the structural formula of lactic acid (CH_3–CHOH–COOH).
- Reaction equation:

CH_3–CHOH–COOH+NaOH→CH3–CHOH–COONa+H_2O

- Calculate the molar concentration of lactic acid in milk from Veq.

Use molar mass M = 90 g/mol.

- Express acidity in Dornic degrees (°D).
- Decide if milk is fresh (≤ 18 °D).
- Determine the Dornic degree of cheese or yogurt.

PRACTICAL WORK N^{O}. 05 :

REDOX TITRATION (MANGANIMETRY)

1. INTRODUCTION

Redox titrations are quantitative analytical techniques based on oxidation–reduction reactions, where electrons are transferred from a reducing agent to an oxidizing agent. Unlike acid-base titrations, these titrations do not require external indicators, because many oxidizing or reducing agents exhibit distinctive color changes.

In this experiment, we study the oxidation of ferrous ions (Fe^{2+}) by the permanganate ion (MnO_4^-) in acidic medium. This type of titration is called manganimetry.

The oxidized form MnO_4^- is purple, while the reduced form Mn^{2+} is colorless. This sharp color transition makes potassium permanganate an excellent self-indicator.

2. OBJECTIVE OF THE EXPERIMENT

To determine the normality and concentration of a ferrous sulfate solution ($FeSO_4 \cdot 7H_2O$) using a standard potassium permanganate solution ($KMnO_4$).

3- FUNDAMENTAL DEFINITIONS

- **Oxidation:** Loss of electrons by a substance.

Example : $Fe \rightarrow Fe^{2+} + 2e^-$

- **Reduction :** Gain of electrons by a substance.

  Example : $Cu^{2+} + 2e^- \rightarrow Cu$

- **Oxidizing Agent (Oxidant):** The species that accepts electrons and is reduced in the process.

  Example: MnO_4^-

- **Reducing Agent (Reductant):** The species that donates electrons and is oxidized in the process.

  Example: Fe^{2+}

- **Redox Couple :** A pair consisting of an oxidized and a reduced form of the same element.

Redox Couple	Oxidized Form	Reduced Form
Fe^{3+} / Fe^{2+}	Fe^{3+}	Fe^{2+}
MnO_4^- / Mn^{2+}	MnO_4^-	Mn^{2+}
Cu^{2+} / Cu	Cu^{2+}	Cu
Ag^+ / Ag	Ag^+	Ag

4- POTASSIUM PERMANGANATE ($KMNO_4$)

- ❖ Appearance: Purple crystalline solid, soluble in water.
- ❖ Chemical nature : Strong oxidizing agent.
- ❖ Role in titrations: Acts both as a reactant and as a self-indicator.
- ❖ Color change at equivalence: From colorless to pale pink when a slight excess of MnO_4^- remains.

4-1- APPLICATIONS :

- ❖ Quantitative analysis (Fe^{2+}, oxalates, hydrogen peroxide).
- ❖ Water treatment (oxidizing iron and manganese in groundwater).
- ❖ Medical field (antiseptic in dilute solutions).
- ❖ Domestic use (removal of mold and fungal stains).

4-2- REACTION MECHANISM OF THE TITRATION

Oxidation half-reaction (iron):

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

Reduction half-reaction (permanganate):

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

Overall balanced reaction:

$$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$$

KEY NOTES FOR STUDENTS

- ❖ **Always work with freshly prepared Fe^{2+} solutions, since they are easily oxidized by air.**
- ❖ **The equivalence point is detected by the first permanent pink color.**
- ❖ **Never use concentrated acids; only dilute H_2SO_4 is suitable.**

5- EXPERIMENTAL PART :

- **Materials and Products :**

Materials	
Burette,	$KMnO_4$ solution (0.1 N).
Pipette,	Ferrous sulfate solution ($FeSO_4\cdot7H_2O$).
Erlenmeyer flask (250 mL	Sulfuric acid (H_2SO_4, 10%).
graduated cylinder,	Distilled water.

EXPERIMENTAL STEPS

1. Rinse all glassware with distilled water, then condition the burette with $KMnO_4$ solution.

2. Fill the burette with $KMnO_4$ solution (N_a = 0.1 N).

3. Using a pipette, transfer 10 mL of $FeSO_4\cdot7H_2O$ solution into a 250 mL Erlenmeyer flask.

4. Add 5 mL of 10% H_2SO_4 and about 50 mL of distilled water.

5. Place the flask on a white sheet of paper to better observe the color change.

6. Perform a rough titration to estimate the equivalence volume.

7. Repeat the titration at least twice for accuracy, recording the exact volume (V_a) of $KMnO_4$ solution consumed.

6-EVALUATION :

- Write the half-reactions and identify the redox couples involved.
- Write the overall redox reaction.
- From the average titrant volume (V_a), calculate:
  - The normality (N_B) of $FeSO_4\cdot7H_2O$.
  - The molar concentration (C_B) of $FeSO_4\cdot7H_2O$.
- Discuss the role of sulfuric acid (H_2SO_4): it provides the acidic medium required for MnO_4^- reduction to Mn^{2+}.
- Identify the limiting reagent of the reaction.
  - Discuss whether H_2SO_4 can be replaced by HCl or H_3PO_4. Explain

PRACTICAL WORK N^{O}. 06 :

REDOX TITRATION (IODOMETRY)

1- INTRODUCTION

Iodometry is a widely used redox titration method that involves iodine (I_2) as an intermediate reagent. Unlike manganimetry, which directly uses a strong oxidant such as potassium permanganate, iodometry is an indirect method. The analyte (the substance being determined) reacts with excess iodide ions (I^-) to liberate iodine (I_2). The liberated iodine is then titrated with a standard solution of sodium thiosulfate ($Na_2S_2O_3$). This reaction is highly specific and provides accurate determinations of many oxidizing agents, such as copper (II), chlorine, and hydrogen peroxide.

2- FUNDAMENTAL DEFINITIONS

• **Oxidation:** Loss of electrons by a chemical species.
Example: $2I^- \rightarrow I_2 + 2e^-$

• **Reduction:** Gain of electrons by a chemical species.
Example: $I_2 + 2e^- \rightarrow 2I^-$

• **Oxidizing Agent (Oxidant):** The species that accepts electrons (e.g., Cu^{2+}, Cl_2, H_2O_2).

• **Reducing Agent (Reductant):** The species that donates electrons (e.g., I^-, $S_2O_3^{2-}$).

• **Redox Couple:** A pair formed by an oxidized and reduced form.
Examples: I_2/I^-, Cu^{2+}/Cu^+, $S_4O_6^{2-}/S_2O_3^{2-}$

3- PRINCIPLE OF IODOMETRY

3-1-. Liberation of iodine :

The oxidizing agent reacts with excess potassium iodide (KI) in acidic medium to liberate iodine (I_2):

Oxidant + $I^- \rightarrow I_2$ + Reduced form

Example :

$2Cu^{2+} + 4I^- \rightarrow 2CuI(s) + I_2$

3-2- Titration of iodine with sodium thiosulfate ($Na_2S_2O_3$):

$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

3-3- End-point detection:

- A starch solution is added as an indicator.

- Iodine forms a deep blue complex with starch.
- At equivalence, the blue color disappears completely when the last traces of iodine are reduced.

4- OBJECTIVE OF THE PRACTICAL WORK

• To determine the concentration of an oxidizing agent (e.g., Cu^{2+}, H_2O_2, or Cl_2 solution) by means of iodometric titration with sodium thiosulfate.

• To understand the indirect role of iodide ions in releasing iodine, which is then titrated.
• To practice using starch as a sensitive indicator for iodine detection.

5- EXPERIMENTAL PART :

- **Materials and Products :**

Materials and Reagents

Materials	
funnel,	Standard sodium thiosulfate solution ($Na_2S_2O_3$).
Burette,	Potassium iodide solution (KI, 10%)
Pipette,	Dilute sulfuric acid (H_2SO_4, 10%).
Erlenmeyer flask (250 mL	Starch indicator solution (freshly prepared).
Beakers and wash bottle	Oxidizing agent to be analyzed ($CuSO_4$ solution).

- **Preparation of the sample:**
  - Pipette a known volume (25 mL) of the oxidizing agent solution into a 250 mL Erlenmeyer flask.
  - Add about 10 mL of KI solution and 5 mL of dilute H_2SO_4.
- **Liberation of iodine:**
  - The oxidant reacts with iodide ions, releasing iodine, which gives a brownish-yellow color to the solution.
- **Titration with sodium thiosulfate:**
  - Fill the burette with $Na_2S_2O_3$ solution.
  - Titrate the iodine until the solution becomes pale yellow.

- **Indicator addition:**

  - Add a few drops of starch solution. The mixture turns deep blue.

- **Final titration:**

  - Continue titration drop by drop until the blue color disappears completely.

  - Note the final volume of $Na_2S_2O_3$ used (V_a).

- Repeat the titration at least twice to obtain concordant results.

6-EVALUATION :

1. **Half-reactions:**

   - Liberation of iodine : $2Cu^{2+} + 4I^- \rightarrow 2CuI + I_2$

   - Reduction of iodine : $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

2. **Global Reaction:**

   $2Cu^{2+} + 4I^- + 2S_2O_3^{2-} \rightarrow 2CuI + 2I^- + S_4O_6^{2-}$

3. **Calculations:**

   - Moles of $Na_2S_2O_3$ used: $n = C \times V$

   - Using stoichiometry, deduce the concentration of the oxidant.

4. Why must starch indicator be added only near the equivalence point?
5. What would happen if HCl was used instead of H_2SO_4?
6. Why must $Na_2S_2O_3$ solutions be freshly standardized?
7. Can iodometry be applied to reducing agents ?

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