

People's Democratic Republic of Algeria
Ministry of Higher Education and Scientific Research



A. MIRA University – BEJAIA

Faculty of Technology

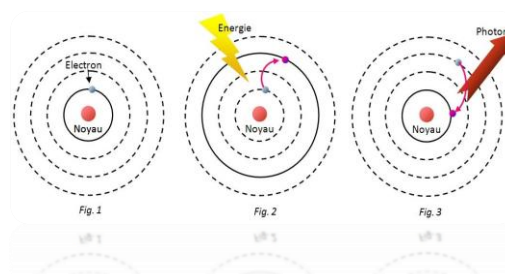
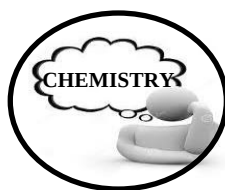
Department of Technology

Common Core Engineering ST (1st year)

Common Core Engineering MT (1st year)



Practical works of Chemistry I



TP N°1: Introduction to Chemistry Experimentation

TP N°2: Solution preparation – Uncertainty calculation – Ionic radius determination

TP N°3: Bohr atomic model – Identification of metal ions by the flame test

TP N°4: Hydrogen ray spectrum and determination of the Rydberg constant

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Academic year 2026/2027

Manipulation n° 1

Introduction to Chemistry Experimentation

I. Introduction

Chemistry lab work is a first experience for most students beginning their university studies. For this time in the laboratory to be successful, certain values must be instilled in the mind of the future chemist (student) in order to carry out the experimental work safely.

Working in a chemistry lab requires handling fragile glassware, heating equipment, complex setups, and potentially hazardous chemicals. It is essential to be aware of the risks associated with each procedure and to be able to respond appropriately in case of an accident.

For the smooth running of chemistry practical sessions, students must ask themselves certain questions: How should we behave in a laboratory? What precautions should we take? What will we be handling? How should we protect ourselves? How should we handle the chemicals?... To answer these questions, a number of rules must be known and applied throughout the practical sessions to ensure one's own safety and that of others.

II. Internal regulations

- Attendance at chemistry practical sessions is mandatory. The instructor must record absences for each session for student evaluation purposes.
- The teacher must monitor attendance and must inform the administration of absences.
- In the event of a justified absence from a practical session, the student is entitled to a make-up session, which will be scheduled by the instructor. If the absence is not justified, a grade of 00/20 will be given, and no make-up session will be offered.
- Three unjustified absences from practical sessions will result in exclusion from the module for the current semester.
- Absences from practical sessions must be justified to the administration within 72 hours of the absence. An official document (proof of absence) will be issued and must be submitted to the instructor within one week. Proofs submitted after this deadline will not be accepted.
- The student must respect the advice and instructions of their instructor.

III. Hygiene and safety rules

Before beginning experimental work in a chemistry laboratory, the student must:

III.1. Have appropriate behavior

- Do not **eat** or **drink** .
- Do not chew **gum** .
- Do not use a **mobile phone** .
- Keep the walkways and exits clear and do not clutter the ground with bags, .
- Do not clutter the workbenches with belongings (binders, pencil cases, etc.).
- Do not play with the equipment.
- Do not taste or smell the chemicals.
- Do not run or move between the lab benches without your teacher's permission.
- Manipulate standing.
- Avoid any rash or impulsive actions or behavior, even if you are in a hurry to finish your lab. **Stay focused** and don't let yourself get distracted.
- Don't arrive late.
- Respect the assignments (laboratory number, group, subgroup, schedule, etc.).

III.2. Wear appropriate clothing

- A long cotton lab coat. Wearing the lab coat is mandatory. Any student who does not bring their lab coat will be sent home by the instructor. The lab coat must remain closed during the manipulation.
- Long trousers and closed shoes to minimize areas of skin exposed in case of splashes.
- Long hair tied back.
- No contact lenses as they can be attacked by volatile solvents.
- If necessary, gloves and safety glasses.

III.3. Take precautions to prevent danger

- Never force glassware when inserting a thermometer into a stopper or a pipette into a pipette filler to avoid cuts from broken glass. Always hold the glass tube near the end to be inserted.
- All handling of flammable products must be carried out carefully and away from flames .
- Never use an unlabeled product and never mix products without first knowing the product of their reaction.
- To avoid contaminating the solution, never take the product directly from the bottle; always use a beaker. Similarly, never return a product or solution that has already been taken to the bottle.
- Never put water onto concentrated acids, but rather the acids onto water.

- Never pipette solutions such as acids, bases, and any toxic or corrosive solutions by mouth. A pipette filler must be used .
- Do not move the equipment.
- Wash your hands before leaving the laboratory.

III.4. Preparing your practical work

- Have the lab manual available from the first session.
- Have a lab notebook.
- Prepare the practical work at home. It is essential to familiarize yourself with the theoretical data and to follow the established procedures for carrying out the experiment.
- Prepare the calculation methods in advance, in case the report is to be carried out on site.
- Carefully follow the steps mentioned on the manual and if in doubt, ask the teacher for advice.
- Apply yourself, observe, and take notes.
- Record the results in the laboratory notebook or on a measurement sheet.
- Write a report of the experiment by answering the questions on the manual.

The student must leave the workbench and glassware clean and tidy. Before leaving the laboratory, each student must return the equipment to its place at the back of the workbench and check that the tap is closed.

IV. Practical session schedule

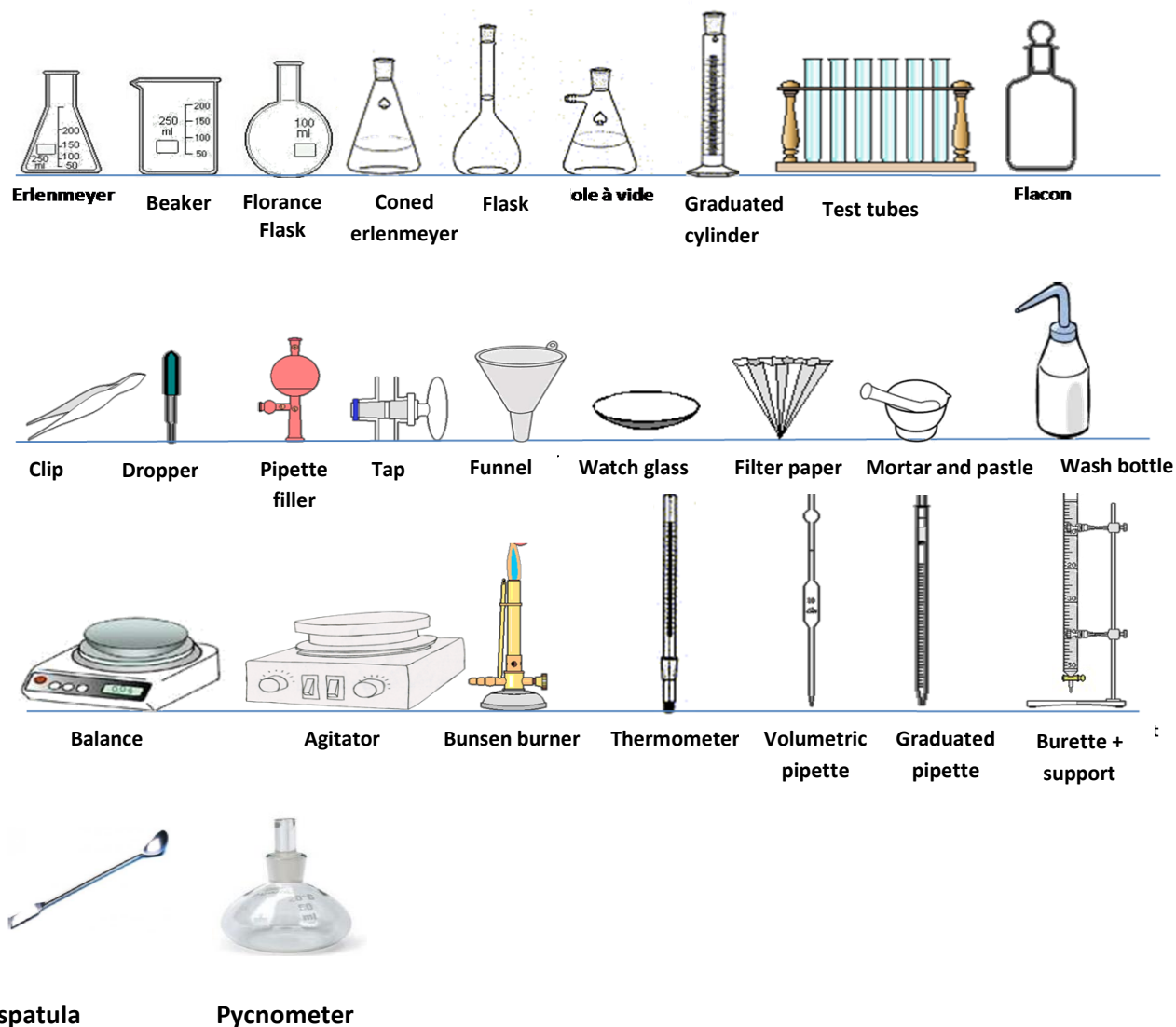
The three hours' practical session will proceed as follows:

- Time will be allocated for practical experiments during which students will be required to follow the procedure precisely. The instructor must supervise, correct, and grade the students' work.
- The rest of the time will be devoted to writing the report, the prototype of which will be given to the student at the end of the experiment.
- Oral questions will be asked of the student during the practical session to better assess and guide them.

V. Chemical hazard labels (pictograms)

GHS Pictogram	Pictogram Name	Hazards Represented
	Exclamation Point	› Irritant to eyes, skin and respiratory tract › Skin sensitizer › Acute toxicity (harmful)
	Health Hazard	› Carcinogen › Reproductive Toxicity › Target organ toxicity › Mutagenicity › Aspiration toxicity › (Effects usually occur over time)
	Skull and Crossbones	› Acute toxicity (severe or fatal) › (Effects can occur shortly after exposure) › (Poisons)
	Corrosive	› Corrosive and damaging to skin and eyes › Chemical burns › Corrosive and damaging to metals
	Flame	› Flammables › Self-reactives › Pyrophorics › Self-heating › Emits Flammable Gas › Organic Peroxides
	Gas Cylinder	› Gases under pressure
	Flame Over Circle	› Oxidizing gases, liquids, or solids (perchlorates) › (Oxidizers are chemicals that facilitate burning or make fires burn hotter and longer.)
	Exploding Bomb	› Explosives › Self-reactives › Organic peroxides › (especially reactive or shock sensitive)
	Environment	› Aquatic toxicity › (not human health)

VI. Laboratory equipment and glassware



Burette: It's a graduated cylindrical glass tube, which ends in a tap and a narrow tube, and fixed in its middle on a support by a Mohr clamp.

Pipette: This is a tube, with or without a reservoir, which allows precise quantities of liquid to be drawn using a bulb. There are two types of pipettes: graduated and volumetric.

Test tube: This is a graduated cylinder which allows for the measurement of very precise volumes, but its precision is low compared to that of the pipette.

Volumetric flask: This is a flask with a narrow neck and a calibration mark. It is used for the preparation or dilution of solutions.

Funnel: It is used to facilitate the introduction of liquids into flasks, burettes.... There are also funnels for solids with a wider lower tube.

Beaker : This is a graduated cylindrical container used to store solutions or samples.... (NB: the beaker is not used to measure precise volumes).

Erlenmeyer: It has a conical shape and is generally used for titrations.

Watch glass: It allows for weighing quantities of products.

Spatula: It is used to take quantities of powdered products from the bottles.

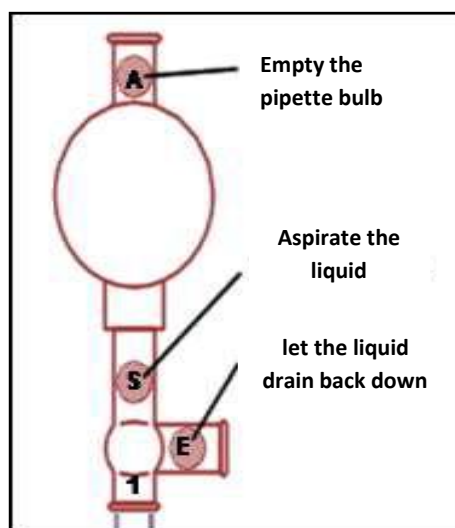
Scale: It allows you to weigh very precise masses of products. There are several types of scales. In our case, we use electronic scales.

VII. Use of certain glassware

VII.1. Pipette

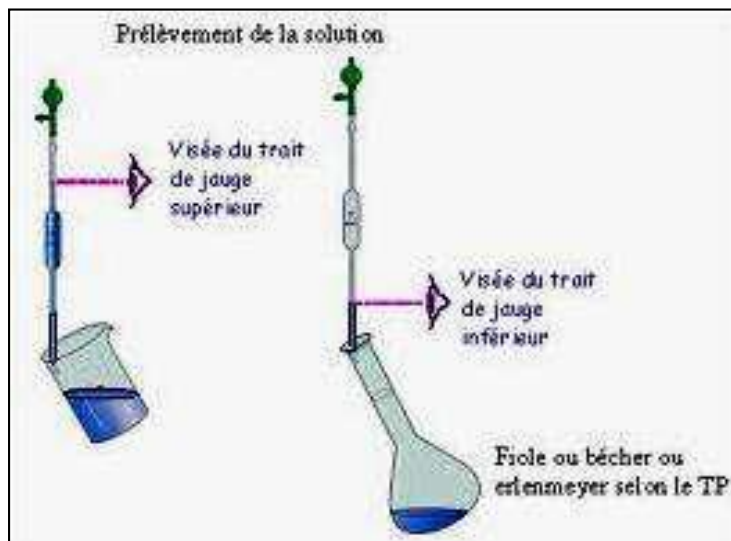
Mouth pipetting is strictly prohibited due to the risk of swallowing a toxic or corrosive solvent. Therefore, a pipette filler , which is attached to the end of the pipette, is used.

In order to create the suction that will allow the liquid to be drawn:



- Squeeze the bulb by pressing on valve "A".
- Push the tip of the pipette lightly into the pipette bulb in the opening provided for this purpose (marked 1 on the diagram).
- Immerse the pipette in the solution and aspirate the liquid by pressing the "S" valve.
- Empty the liquid into the beaker by pressing valve "E". To prevent the liquid from pouring too quickly, hold the beaker and pipette at a 45° angle by placing the tip of the pipette against the side of the beaker while keeping it vertical.

Reading : the line should be drawn at eye level.



Cleaning: Take care not to introduce any liquid inside the bulb. If liquid accidentally gets inside, rinse thoroughly with water and dry as best as possible with absorbent paper.

Accuracy of volumetric pipettes : Pipette accuracy is an important criterion for purchase and use. For dosing, only volumetric pipettes can be used, as graduated pipettes are not precise enough.

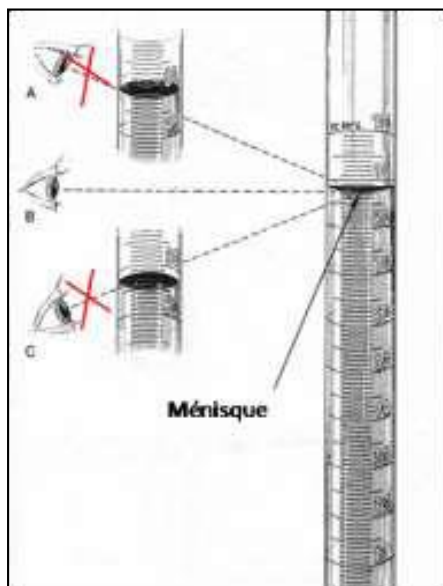
VII.2. Burette

A burette is a graduated and calibrated glass tube with a stopcock used to slowly add reagent during a titration. It is available in several sizes, with 25 mL and 50 mL being the most common. Graduations are generally accurate to 0.05 mL but can also be as precise as 0.1 mL .

Instructions for use:

- Pour a small amount of reagent into the burette to rinse it.
- After rinsing, fill it completely and open the tap to eliminate any air bubbles trapped at the tap.
- Refill it and zero it by slowly emptying the excess titrant solution.
- During the titration, pour volumes in 1 mL increments , tightening the volumes at the equivalence point until a drop- by-drop method is used to obtain sufficient accuracy.
- After dosing, empty the burette completely and rinse with distilled water. When the burette is not in use, refill it with distilled water.

Reading: The volume reading is taken at the bottom of the meniscus, keeping the eye aligned with the graduation to avoid parallax error (see A & C). The bottom of the meniscus must coincide with the graduation. The eye remains aligned with the graduation during the reading (see B).



Maintenance : The tap can be made of glass or Teflon. For those with a glass tap, it is advisable to lubricate it with silicone grease (used for lapping, for example) from time to time to eliminate any risk of seizing.

Manipulation 2

Solution preparation – Uncertainty calculation – Ionic radius determination

I. Introduction

Engineering sciences are based on experiments. These experiments are then compared with theories intended to describe, in a fair and objective manner, what is observed in the real world. However, both approaches, experimental and theoretical, are obviously accompanied by the notion of doubt, a fundamental concept for a future engineer.

It is more than obvious that a physical or chemical quantity cannot be measured without being completely free from error or uncertainty. Yet, every result is of interest provided that it is expressed with its limits of uncertainty.

This practical work will enable the student to calculate the uncertainties involved in experimental measurements, namely the measurement of the concentration of solutions prepared in different ways and the determination of the ionic radius, in order to present their results more accurately.

II. Theoretical Background

II.1. Concepts of solutions and concentrations

Solute: This is the substance, matter, or chemical species dissolved in a solvent.

Solvent: This is the liquid in which the solute is dissolved. The solvent can be distilled water, ethanol, etc.

Solution : This is the name for a homogeneous mixture of solute and solvent. When the solvent is water, the solution is said to be **aqueous** .

Molar mass (M): This is the mass of one mole of a compound expressed in g.mol^{-1} .

Amount of substance (n): This is the number of moles that this sample contains: $n = \frac{m}{M}$ (moles).

Mass density (ρ): This is the mass ratio corresponding to a volume V of a substance: $\rho = \frac{m}{V}$ ($\frac{\text{Kg}}{\text{m}^3}$).

Density (d): This is the ratio of the mass density of a substance to the mass density of a reference: $d = \frac{\rho_{\text{substance}}}{\rho_{\text{ref}}}$.

Molar concentration (Molarity, C): This is the amount of solute present per liter of solution:

$$C = \frac{n_{\text{solute}}}{V_{\text{solution}}} \text{ (mol.l}^{-1}\text{)}.$$

Mass concentration (C_M): This is the ratio between the mass of solute and the total volume of the solution: $C_M = \frac{m_{\text{solute}}}{V_{\text{solution}}} \text{ (gl}^{-1}\text{)}.$

Normal concentration or normality (N): This is the number of gram equivalents of solute contained in one liter of solution: $N = \frac{n_{\text{eq.g}}}{V} \text{ (eq.gl}^{-1}\text{)}.$

Mass fraction (x_m): This is the ratio between the mass of the solute and the mass of the solution: $\bar{x}_m = \frac{m_{\text{solute}}}{m_{\text{solution}}}$

Mole fraction (x): This is the ratio between the number of moles of the solute and the number of moles of the solution: $x = \frac{n_{\text{solute}}}{n_{\text{solution}}}$

Dilution: Addition of solvent to a solution to decrease its concentration.

$$n_{\text{mother}} = n_{\text{daughter}} \rightarrow C_m V_m = C_f V_f$$

II.2. Concepts on the calculation of measurement error:

II.2.1. Uncertainty in a direct measurement: The error can be expressed as:

1. Error : This is the absolute value of the difference between the true value X_T and the measured value X_m . Since the true value X_T is unknown, the absolute error is also unknown.

$$\text{Error} = |X_T - X_m| = \text{unknown}$$

Absolute uncertainty = upper limit of absolute error = ΔX

2. Relative error: is an unknown value which is the ratio of the absolute error to the measured value.

$$\text{Relative error} = \frac{\text{Absolute error}}{\text{Measured error}} = \frac{X_T - X_m}{X_m} = \text{unknown}$$

Relative uncertainty is the upper limit of relative error. It is the quotient of the absolute uncertainty ΔX by the measured value X_m .

$$\text{Relative uncertainty} = \frac{\text{Absolute uncertainty}}{\text{Measured uncertainty}} = \frac{\Delta X}{X_m} = \text{unknown}$$

II. 2.2. Uncertainty on a calculated quantity: the following important cases are considered:

1. **Absolute uncertainty about a sum or difference:** when the composite quantity consists only of sums or differences:

$y = x_1 \pm x_2 \pm x_3$, therefore the absolute uncertainty is:

$$\Delta y = \Delta x_1 + \Delta x_2 + \Delta x_3 + \dots$$

In a sum or a difference, absolute errors are added together.

2. **Uncertainty about a product or quotient:** when the composite quantity consists only of products or quotients:

Consider the relationship between y and x:

a) $y = x_1 * x_2$ then $\ln y = \ln x_1 + \ln x_2$ and $\frac{dy}{y} = \frac{dx_1}{x_1} + \frac{dx_2}{x_2}$

The relative uncertainty is therefore: $\frac{\Delta y}{y} = \frac{\Delta x_1}{x_1} + \frac{\Delta x_2}{x_2}$

b) $y = \frac{x_1}{x_2}$ then $\ln y = \ln x_1 - \ln x_2$ and $\frac{dy}{y} = \frac{dx_1}{x_1} - \frac{dx_2}{x_2}$

The relative uncertainty is therefore: $\frac{\Delta y}{y} = \frac{\Delta x_1}{x_1} + \frac{\Delta x_2}{x_2}$

c) if the relationship between y and x is: $y = \frac{x_1 * x_2}{x_3}$

The relative uncertainty is therefore: $\frac{\Delta y}{y} = \frac{\Delta x_1}{x_1} + \frac{\Delta x_2}{x_2} + \frac{\Delta x_3}{x_3}$

3. **Uncertainty in a power product:** when the composite quantity consists only of a product of powers:

$$y = x_1^a * x_2^b * x_3^c \quad \text{then } \ln y = a * \ln x_1 + b * \ln x_2 + c * \ln x_3 \text{ and } \frac{dy}{y} = \frac{dx_1}{x_1} + \frac{dx_2}{x_2} + \frac{dx_3}{x_3}$$

$$\frac{\Delta y}{y} = a \frac{\Delta x_1}{x_1} + b \frac{\Delta x_2}{x_2} + c \frac{\Delta x_3}{x_3}$$

II.2.3. Error classification

Two main types of errors can be distinguished, which can affect any measurement: systematic errors and accidental or fortuitous errors.

Manipulation 2 : Solution preparation – Uncertainty calculation – Ionic radius determination

2026 - 2027

a . Accidental or Fortuitous Errors: Accidental errors cannot, in principle, be avoided. Their causes lie within the experimenter himself : the sureness with which the hand handles an instrument, the accuracy with which the eye observes. It is the task of every observer to be aware of accidental measurement errors, to keep them as small as possible, and to estimate or calculate their influence on the result obtained.

b. Systematic errors: These are linked to a defect in materials (glassware) or the measurement method. They include, in particular, errors in the construction and calibration of the measuring equipment used. They are repetitive and consistent. Systematic errors must be eliminated.

Uncertainties regarding the main measuring instruments

Volumetric pipettes		Graduated pipettes		Volumetric flasks		Graduated test tubes	
Volume (ml)	Uncertainty (ml)	Volume (ml)	Uncertainty (ml)	Volume (ml)	Uncertainty (ml)	Volume (ml)	Uncertainty (ml)
1	± 0.006	0.2	± 0.008	10	± 0.02	10	± 0.1
2	± 0.006	0.5	± 0.01	25	± 0.05	25	± 0.3
5	± 0.01	1	± 0.01	50	± 0.06	50	± 0.4
10	± 0.02	2	± 0.02	100	± 0.08	100	± 0.6
25	± 0.03	5	± 0.02	200	± 0.1	250	± 1
50	± 0.05	10	± 0.06	250	± 0.1	500	± 3
100	± 0.08	25	± 0.1	500	± 0.2		

III. Purpose of the practical work

- Prepare a solution by dissolving a chemical species and by dilution.
- Measuring the density of a solution.
- Determining the average radius of an ion, which is nothing more than an atom that has lost or gained electrons.
- Learn to calculate uncertainties on the different measurements and quantities used in chemistry lab experiments.

IV. Materials and products used

- Graduated cylinder of 50 ml;
- Volumetric flask of 100 ml;
- Graduated pipette of 5 ml;

- Beaker of 50 ml;
- pycnometer of 10 ml;
- Watch glass;
- Funnel;
- Pissette;
- Distilled water;
- Sucrose; Sodium chloride;
- Concentrated commercial solution of acetic acid;
- Electronic scale, if possible more precise than 0.1g.

V. Manipulations

Experiment 1: Preparation of solution by dissolution

- Weigh a mass m of sucrose $C_{12}H_{22}O_{11}$ using an electronic balance and a dish or a watch glass ($m = 1g; 1.5g; 2g$)
- Introduce the solid into a 100 ml volumetric flask using a funnel.
- Rinse the container used and the funnel with a wash bottle of distilled water.
- Fill the volumetric flask approximately $\frac{3}{4}$ full with distilled water.
- Shake to accelerate dissolution and homogenize the solution.
- Top up with distilled water to the gauge line.
- Solution S_1 is ready.

Experiment 2: Preparation of solution by dilution

- Calculate the volume of a concentrated commercial solution of acetic acid of density $d = 1.02$, containing 30% by mass of CH_3COOH , required to prepare 100 ml solution of CH_3COOH of concentration $C = 0.1 - 0.2 - 0.3 \text{ mol.L}^{-1}$.
- Put a sufficient quantity of commercial acetic acid into a beaker.
- Measure out the previously calculated volume using a pipette or a graduated cylinder.
- Put this volume into a volumetric flask containing a little distilled water, then fill with water up to the calibration mark.
- Solution S_2 is ready.

Experiment 3: Experimental determination of the mass density and density of a solution

- Using a balance, weigh a 10 ml pycnometer. This mass will be noted as m_1 .

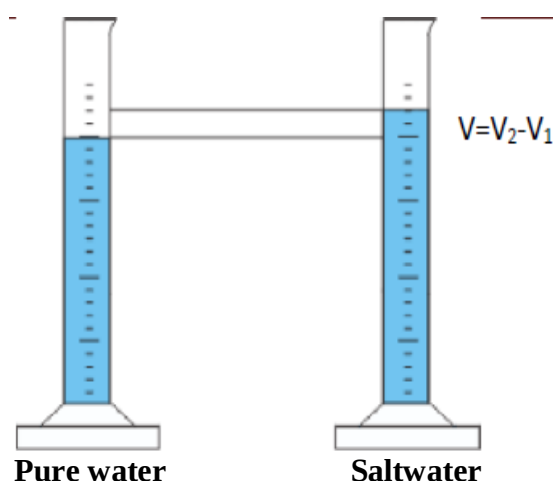
Manipulation 2 : Solution preparation – Uncertainty calculation – Ionic radius determination

2026 - 2027

- Fill this pycnometer with the previously prepared solution S_1 .
- Weigh the assembly: pycnometer + 10ml of solution. This mass will be noted as $m_{\text{pycnometer+solution}}$.
- Repeat the same procedure for solution S_2 .

Experiment 4: Experimental determination of the ionic radius

Salt is poured into a certain volume of water; when it is completely dissolved, the volume of salt water is greater than the initial volume. The ions released by the dissolution of the salt crystal (sodium chloride) thus occupy space, which is measurable by the increase in the volume of the liquid:



- Weigh a mass of salt ($m_{\text{NaCl}} = 5\text{g}$).
- Accurately measure a certain volume of fresh water ($V_1 = 30\text{ml}$) in a 50ml test tube.
- Dissolve the NaCl salt completely in the test tube to obtain ions (Na^+ and Cl^-); making sure to eliminate all air bubbles.
- Measure the volume V_2 of the solution after adding and dissolving the salt and deduce the change in volume ΔV corresponding to the volume of one ion V_{ion} .

VI. Questions:

1. Calculate the value of the molar and mass concentration of the sugar solution.
2. Calculate the molality and mole fractions x_i of the solute and solvent of the prepared solution S_1 .
3. Calculate the concentration of the commercial acetic acid solution with 30% by mass.

Manipulation 2 : Solution preparation – Uncertainty calculation – Ionic radius determination

2026 - 2027

4. Indicate the method and the value of the volume of CH_3COOH that must be taken for the preparation of the acetic acid solution S_2 .
5. Determine the mass density and the density of solutions S_1 and S_2 .
6. Calculate the relative error made on the concentration, mass density and density of solutions S_1 and S_2 .
7. Give the ionization reaction of the salt NaCl .
8. Calculate:
 - a. The number of moles of salt.
 - b. The number of dissolved ions, knowing that a NaCl molecule releases two ions (Na^+ and Cl^-).
 - c. The volume of an ion

$$V_{\text{ion}} = \frac{V}{N_{\text{ion}}}$$

- d. Assuming that the Na^+ and Cl^- ions are of the same order of magnitude and spherical shape, calculate the radius of these two ions.

$$\text{We are given: } N_{\text{ion}} = \frac{4}{3} \pi r^3$$

- e. Calculate the value of the relative (%) and absolute (Δ) uncertainty committed on the measurement of the ionic radius r_{ion} .
- f. Give the ionic radius with its uncertainty, written in the form $r_{\text{ion}} \pm \Delta r_{\text{ion}}$.
- g. Compare the calculated value to the value measured by X-ray diffraction, which is 0.14 nm. Are the values the same? If not, explain the difference.

Manipulation 3

Bohr atomic model – Identification of metal ions by the flame test

I. Introduction

An atom consists of a central, spherical nucleus. This nucleus is composed solely of protons and neutrons. The nucleus is held together by the strong interaction of a cloud of electrons orbiting around it. This electron cloud is represented by electron layers (K, L, M...) and sublayers (s, p, d...). Each layer represents a specific energy level in the ground state.

According to quantum theory, developed by Planck and then by Einstein, light is made up of a collection of massless particles called photons, each carrying a quantum of energy equal to:

$$\Delta E = h \nu = h \frac{c}{\lambda}$$

With :

h: Planck constant (6.63×10^{-34} Js)

E: energy of the photon in Joules

ν : time frequency (s^{-1})

λ : wavelength (m)

c: speed of light (3×10^8 ms⁻¹)

It was after studying the light emitted by certain heated bodies that N. Bohr's research team, in 1913, improved Rutherford's atomic model by introducing the model of electron layers. Bohr's model states that electrons exist only at certain allowed energy levels. When an atom heats up, it absorbs energy. Electrons transition from their ground state (lowest energy state) to an excited state, and then the atom returns to its stable state, releasing its energy in the form of a photon. This emitted photon is defined by a quantized, precise, and specific wavelength λ . In the visible spectrum, energy is expressed as colored light characteristic of a specific wavelength. This wavelength depends on the difference between the two energy levels.

Each element has its own characteristic set of energy levels. Each metal produces a characteristic flame emission spectrum, which translates into a flame color. Studying these emission spectra allows for the identification of metallic elements.

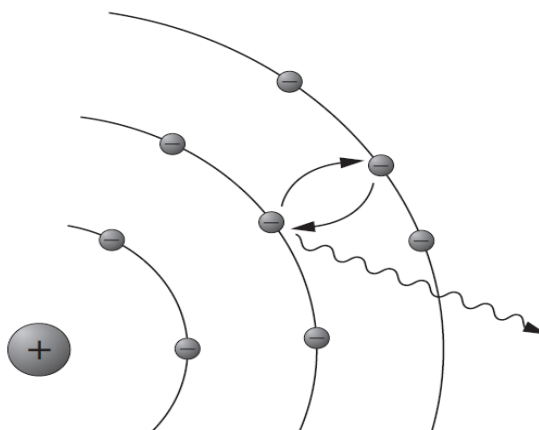
II. Flame test

II.1. Definition

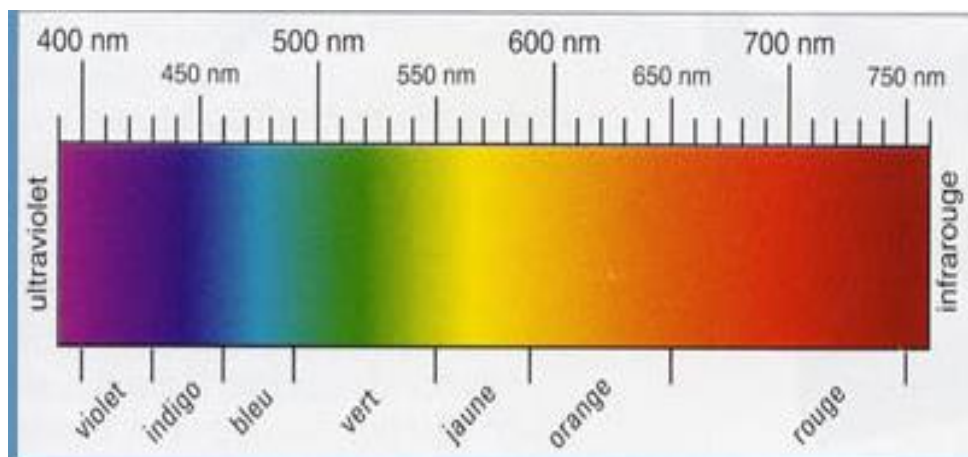
A flame test is a method used in chemistry to detect the presence of certain metal ions, based on the emission spectra of each element. Generally, the color of the flame also depends on the

temperature. The test involves introducing a sample of the element or compound to be analyzed into a non-luminous, high-temperature flame and then observing the resulting color.

When a metallic element in the form of a salt is heated in a flame, or when a high-voltage electrical source is applied to a gas, the atoms absorb energy. An atom then possesses more energy than in its fundamental state; it is said to be "excited." Thanks to the absorbed energy, an electron can move to a higher energy level. In other words, the electron is propelled to a layer farther from the nucleus. However, the excited atom is unstable and quickly returns to its fundamental state. The electron falls back to a layer closer to the nucleus. During the electron's fall from an excited state to a lower state, the absorbed energy is released as light energy. Depending on the amount of energy released, the light will have a different color, which demonstrates the existence of different energy levels.



The color of the flame is not related to the color of the metallic salt solution. For example, copper sulfate solution is blue, but its flame color is green.



couleur	longueur d'onde en nanomètre	Metallic cation	Flam color
rouge	~ 625-740	<u>Copper</u> (Cu^{2+})	Bluish green
orange	~ 590-625	<u>Baryum</u> (Ba^{2+})	Yellowish green
jaune	~ 565-590	<u>Strontium</u> (Sr^{2+})	Bright red
vert	~ 520-565	<u>Calcium</u> (Ca^{2+})	Orange red
bleu	~ 446-520	<u>Potassium</u> (K^{+})	Lilas
violet	~ 380-446	<u>Lithium</u> (Li^{+})	Crimson red
		<u>Sodium</u> (Na^{+})	Orange yellow



II.2. Relationship between the light emitted by heated bodies and the energy levels (n) of electrons

To explain the color emissions of metals and their salts, we must think at the atomic level by studying two stages in the behavior of atoms:

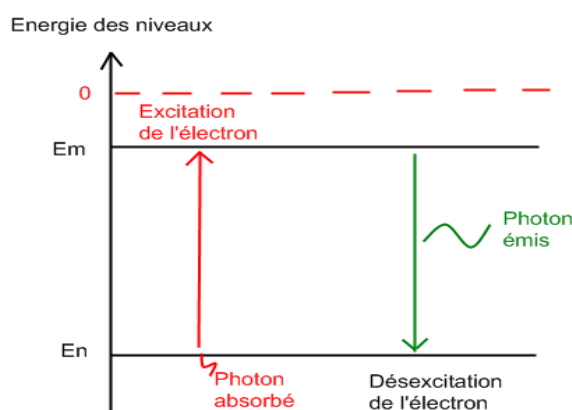
- Energy absorption
- Energy recovery

❖ Energy absorption

Before the experiment, and therefore before being heated, the atoms are in a stable state called the "fundamental state." During the experiment, the flame provides heat energy. This energy is absorbed by the electrons. The energy level of the electrons then increases, moving from the stable fundamental energy level (E_n) to a higher energy level (E_m). At this energy level, the electrons are unstable. They are said to be "excited." This is the absorption of energy.

❖ Energy emission

After heating and vaporization, the atoms rising towards the tip of the flame find themselves in a cooler region. They then release the energy they had absorbed in the form of light energy.

**III. Purpose of the practical work**

This lab aims to demonstrate that electrons in atoms are located in specific energy levels, and not randomly. When excited by heat, electrons transition from stable to unstable (higher energy) levels. Upon de-excitation, they return to their original energy level and emit light of a precise wavelength (color). In this lab, we will identify unknown metal ions using the flame method, by comparing them to the colors of flames produced by solutions of known metal salts.

IV. Materials and Products

- Bunsen burner (or lighter)
- Test tubes containing the solutions of the salts to be studied
- Wooden rods or glass rods
- Ethanol or methanol
- Cotton

-Metallic salt solutions:

Sodium chloride (NaCl)

Barium chloride (BaCl_2)

Potassium chloride (KCl)

Lithium chloride (LiCl)

Calcium chloride (CaCl_2)

Copper sulfate (CuCl_2)

V. Operating Procedure

Prepare the metal salt solutions by adding the equivalent of 2 teaspoons of powder to test tubes, then adding water to dissolve the salt. Wash the spatula between each powder to avoid mixing. Dip a cotton-wrapped wooden rod into each tube. Pass the wooden rods or glass rods soaked in the salt solutions through the flame of a lighter (the blue part of the flame) separately over the flame. (To avoid burning the wood, move the wooden rod through the flame rather than holding it upright.) Record the flame color in the data table. Repeat each experiment at least twice to ensure the results are accurate.



VI. Questions

- 1- Is there an ion among the studied ions that could be a hydrogenoid?
- 2- If so, which one? Give its formula.
- 3- Give the definition of a hydrogenoid.
- 4- Why is it interesting to work with this type of ion?
- 5- Define a flame test by explaining the phenomenon involved.

- 6- Provide the flame color for each sample.
- 7- Give the approximate wavelengths based on the colors given in the figure glued to your workbench.
- 8- Identify the metal ions contained in your solutions based on the flame colors of the ions in the figure glued to your bench.
- 9- Place the observed colours in ascending order of wavelength.
- 10- Place the observed colors in ascending order of energy. Which elements correspond to the lowest and highest energy?
- 11- What is the relationship between energy, frequency, and wavelength?
- 12- How are electrons excited during this experiment?
- 13- Give the emission spectrum of light by the electrons.
- 14- Which particles in the solutions would be responsible for the production of colored light?
- 15- Why do different chemicals emit different colors of light, especially when heated?

Manipulation 4

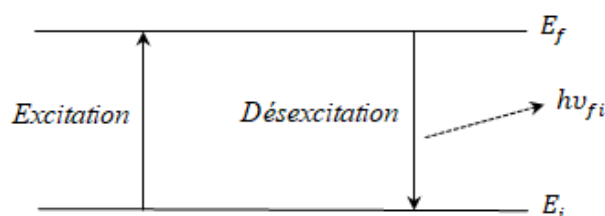
Hydrogen ray spectrum and determination of the Rydberg constant

I. Theoretical background

I.1. Energy level of an atom

The energy of an atomic system is quantized, meaning that the system can only exist in specific (stationary) states where its energy has well-defined values called energy levels. The same is true for its orbital angular momentum and its magnetic moment. The lowest energy state is the most stable (fundamental state).

In the case of an isolated atom (no interaction with the universe), the atom is in a "stable" state (stationary state). If it is disturbed by an external cause (discharge, electron collisions, radiation, etc.), it can transition from an initial energy state E_i to a final energy state E_f by absorbing energy $E_{i \rightarrow f} = E_f - E_i$. The atom is then in an excited state with a variable and finite lifetime. During de-excitation, the return of the atom to state E_i causes the emission of a photon whose energy is: $h\nu = E_f - E_i$:



I.2. Ray Spectrum

The emission spectrum is characteristic of the material and can serve as a fingerprint to identify the gas. Ray spectra therefore serve as a key to understanding the structure of the atom: any theory of atomic structure must explain why atoms emit only light of discrete wavelengths and must be able to predict these frequencies.

Atoms emitting light in a luminous gas produce spectra made up of numerous spectral lines that are separated from each other, even though they may appear in large numbers in certain areas. These rays are therefore characteristic of the chemical element because each one corresponds to a transition between two distinct energy levels within the atom.

Hydrogen is the simplest atom and also has the most elementary spectrum. The spectral rays of hydrogen have been explained by Bohr's atomic model. However, some time before, the rays

in the hydrogen spectrum could be determined by an empirical formula that included the Rydberg constant.

Electrons undergoing thermal agitation in an incandescent light source emit electromagnetic radiation (light) composed of different wavelengths, thus producing a spectrum. By analyzing the light emitted from an excited gas, a vaporized liquid, or a solid, spectral lines are observed.

In the visible range, the emission spectrum of atomic hydrogen shows four lines H_{α} , H_{β} , H_{γ} and H_{δ} which continue in a complete series in the ultraviolet (Figure 1).

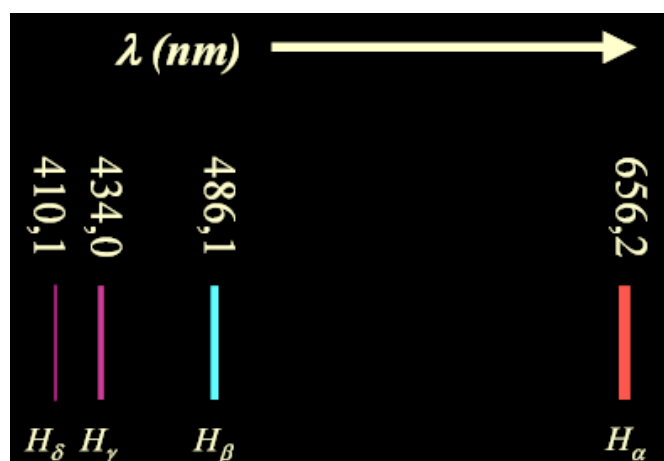


Figure 1: Visible spectrum of hydrogen.

Modern physics explains these spectra in terms of photons of light with discrete wavelengths emitted during transitions between two energy levels. Each substance possesses a set of spectral rays with well-defined wavelengths. In 1885, J.J. Balmer (1825–1898) showed that the four lines visible in the spectrum of hydrogen (whose wavelengths are, experimentally, in good agreement with those obtained theoretically) obeyed the following empirical formula:

$$\frac{1}{\lambda} = R \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \text{ with } n = 3, 4, 5, \dots$$

Where R is the Rydberg constant, with a value of $R = 1.097 \times 10^7 \text{ m}^{-1}$. The spectrum of hydrogen has an interesting theoretical characteristic: having only one proton and one electron, it is the simplest atom. Neils Bohr (1885–1962) developed an atomic theory for hydrogen, arguing that the spectral lines resulted from the transition of the electron from a higher to a lower energy level. The wavelength of the spectral lines could be given by the theoretical relation:

$$\lambda = \frac{h c}{\Delta E}$$

Or: $\Delta E = E_1 \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \rightarrow \Delta E = 13,6 \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \text{ in eV}$

ΔE : energy difference between the initial level n_i and the final level n_f ;

h : Planck's constant ($h = 6.62 \times 10^{-34} \text{ Js}$);

c : speed of light in a vacuum ($c = 3 \times 10^8 \text{ m/s}$);

E_1 : ground state energy.

The values $n = 1, 2, 3, 4$ are called principal quantum numbers.

II. Purpose of the practical work

- Record the ray spectrum of hydrogen.
- Determine the wavelengths and frequencies of the spectral lines H_α , H_β , H_γ and H_δ of the Balmer series of hydrogen.
- Calculate the energy levels of the hydrogen atom.
- Calculate Rydberg's constant.

III. Equipment used

- 1- Digital LD spectrometer;
- 2- Power supply for spectral tubes;
- 3- Hydrogen spectral lamp;
- 4- Clamping base;
- 5- Fiber optics;
- 6- Support by rod and terminal for optical fiber;
- 7- Computer.



Figure 2: Experimental setup.

IV. Manipulation

- Check the setup including the digital spectrometer, the power supply for spectral tubes containing the integrated hydrogen lamp and the computer.
- Ensure that the room is not too brightly lit to minimize the influence of light.
- Place the optical fiber as close as possible to the aperture of the spectral tubes' feed in order to intensify the acquisition signal.
- Turn on the computer.
- Go to the " Spectro " file and click on the " Esaostudio " application to launch the software which will then run automatically.

In the software, the user interface window is displayed in "Emission" working mode and measurement values are immediately recorded.

- Uncheck the exposure time.
- Choose the maximum exposure time.
- Sensitivity condition" box checked .
- Turn on the power supply for the spectral tubes and wait 10 to 15 minutes (preheating) for a good acquisition before recording the spectrum in "Emission" mode.
- Conclude the recording of the emission spectrum by clicking on the "Run/Pause" button as soon as the spectral lines are clearly visible.
- Give the file a name by entering it in the "Name" field.
- Save the spectrum by clicking the "Save" button.
- Read the wavelengths of the spectral lines and note them down . To read them, move the mouse pointer over the spectral lines. If necessary, select another tool by going to the "Tools" bar and then " Coordinates ".

V. Questions:

1. How many rays are observed? Name them.
2. Why can't we clearly observe the H_{δ} line?
3. As $n \rightarrow \infty$, the wavelengths of the Balmer series approach a minimum or limiting wavelength. Calculate the limiting wavelength for the Balmer series.
4. Why can only four spectral lines be observed in the Balmer series? (Transitions for $n > 6$ also exist). Justify your answer mathematically.
5. What is the spectral range of the hydrogen emission spectrum?

6. Complete table 1.

n	$1/n^2$	Color	λ (nm)	ν (Hz)
3				
4				
5				
6				

7. Plot the graph $\nu = f(1/n^2)$.

8. From the recorded measurements, deduce the Rydberg constant graphically.

9. Calculate the relative error made on the measurement of the Rydberg constant knowing that the theoretical value is 109737 cm^{-1} .

10. Calculate the energy of the levels of the hydrogen atom involved in the Balmer series in eV and in J.

11. Calculate the energy variation (in J and eV) of the photons corresponding to the observed lines ($3 \rightarrow 2$, $4 \rightarrow 2$, $5 \rightarrow 2$ and $6 \rightarrow 2$).

12. Plot $\Delta E = f(1/n^2)$. Deduce the energy E_1 of the ground state.