CHEMISTRY EXPERIMENTS USER GUIDE

FOR

LOWER SECONDARY (S1-S3)

FOREWORD

Dear teacher,

Rwanda Basic Education Board (REB) is honoured to present Chemistry experiments user guide for lower secondary (S1-S3). This book will serve as a guide to competence-based teaching and learning to ensure consistency and coherence in the learning of Chemistry.

In this book, special attention is paid to experiments that facilitate the learning process in which students can manipulate concrete apparatuses and chemicals to carry out appropriate experiments, develop ideas, and make adequate interpretations and conclusions during activities performed individually or in pairs/ small groups.

In competence-based curriculum, experiments open students' minds and provide them with the opportunities to interact with the world, use available tools, collect data and effectively model real life problems.

For efficiency use of this user guide book, your role as a teacher is to:

- Plan your lessons and prepare appropriate teaching materials (chemicals and reagents),
- Organize group discussions for students considering the importance of social constructivism,
- Engage students through active learning methods,
- Provide supervised opportunities for students to develop different competences by giving tasks which enhance critical thinking, problem solving, research, creativity and innovation, communication, and cooperation,
- Support and facilitate the learning process by valuing students' contributions in the practical activities,
- Guide students towards the harmonization of their findings,
- Encourage individual, peer and group evaluation of the work done and use appropriate competence-based assessment approaches and methods.

To facilitate you in your teaching activities, the content of this booklet is self-explanatory so that you can easily use it. It is divided in 3 parts:

The part I explains the structure of this book and gives you the general introduction on laboratory experiments.

The part II gives the list of apparatuses and chemicals needed to perform experiments in the booklet of chemistry.

The part III details the setup of experiments, the procedures to be followed when performing experiments, interpretations of results and conclusions.

I wish to sincerely extend my appreciation to the people who contributed towards the development of this guide, REB particularly those from Curriculum, Teaching and Learning Resources Department and SPIU staff who organized the whole process from its inception. Special appreciation goes also to teachers and independent experts in education and IEE staff who supported the exercise throughout. Any comment or contribution would be welcome for the improvement of this booklet for next versions.

Dr. MBARUSHIMANA Nelson

Director General, REB

ACKNOWLEDGEMENT

I wish to express my appreciation to all the people who played a major role in the development of Chemistry experiments user guide for lower secondary. It would not have been successful without their active participation.

Special thanks are given to IEE, AIMS, independent people, teachers, illustrators, designers and all other individuals whose efforts in one way or the other contributed to the success of the development of this user guide.

I owe gratitude to the Rwanda Basic Education Board staff particularly those from Curriculum, Teaching and Learning Resources Department who were involved in the whole process of the development work.

Finally, my word of gratitude goes to REB-SPIU under RQBE Project for their usual support in terms of human and financial resources towards improving the quality of education.

Joan MURUNGI

Head of CTLR Department

LIST OF ACRONYMS

CBC	:	Competence-based curriculum
ICT	:	Information Communication Technology
Lab	:	Laboratory
STEM	:	Science, Technology, Engineering and Mathematics
CPD ITMS	:	Continuous Professional Development Certificate in Innovative Teaching Mathematics and Science
IBL	:	Inquiry Based Learning
KBC	:	Knowledge Based Curriculum
SET	:	Science and Elementary Technology
UR-CE	:	University of Rwanda- College of Education
CTLR	:	Curriculum, Teaching and Learning Resources

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PART 1: GENERAL INTRODUCTION

1. Laboratory experiments in the Competence Based Curriculum

A competence-based curriculum (CBC) focuses on what learners can do and apply in different situations by developing skills, attitudes, and values in addition to knowledge and understanding. This learning process is learner-focused, where a learner is engaged in active and participatory learning activities, and learners finally build new knowledge from prior knowledge. Since 2015, the Rwanda education system has changed from KBC to CBC for preparing students that meet the national and international job market requirements and job creation. Therefore, implementing the CBC education system necessitates qualitative laboratory practical works for mathematics and science as more highlighted aspects.

In addressing this necessity, laboratory experiments play a major role. A student is motivated to learn chemistry by getting involved in handling various concrete manipulations in various experiments.

For learning chemistry concepts through the above-mentioned approach, Chemistry kits composed by chemicals and apparatuses for ordinary level have been distributed into schools. The kits include various items along with a manual for performing experiments.

The kit broadly covers the experiments that are proposed in the syllabus. The kit has the following advantages:

- Availability of necessary and common materials at one place,
- Multipurpose use of items,
- Economy of time in doing the activities,
- Portability from one place to another,
- Provision for teacher's innovation,
- Low-cost material and use of indigenous resources.

Apart from the kit, the user guide for laboratory and practical activities to be used by teachers was developed. This laboratory experiment user guide is designed to help mathematics and science teachers to perform high-quality lab experiments for mathematics and science. This user guide structure induces learner's interest, achievement, and motivation through the qualitative mathematics and science lab experiments offered by their teachers and will finally lead to the targeted goals of the CBC education system, particularly in the field of mathematics and science.

In CBC, learners hand-on the materials and reveal the theory behind the experiment done. Here, experiments are done inductively, where experiments serve as an insight towards revealing the theory. Thus, the experiment starts, and theory is produced from the results of the experiment.

2. Type of laboratory experiments

The goal of the practical work defines the type of practical work and how it is organized. Therefore, before doing practical work, it is important to have a clear idea of the objective.

The three types of practical work that correspond with its three main goals are:

a) **Equipment-based practical work:** the goal is for students to learn to handle scientific equipment like using a microscope, doing titrations, making an electric circuit, etc.

b) **Concept-based practical work:** learning new concepts.

c) **Inquiry-based practical work:** learning process skills. Examples of process skills are defining the problem and good research question(s), installing an experimental setup, observing, measuring, processing data in tables and graphs, identifying conclusions, defining limitations of the experiment etc.

Note:

- To learn the new concept by practical work, the lesson should start with the practical work, and the theory can be explained by the teacher afterward (explore – explain).
- Starting by teaching the theory and then doing the practical work to prove what they have learned is demotivating and offers little added value for student learning.

- Try to avoid complex arrangements or procedures. Use simple equipment or handling skills to make it not too complicated and keep the focus on learning the new concept.
- If this is not possible and is necessary to use new equipment or handling skills, then first exercise these skills before starting the concept-based practical work experiments.
- The experiments should be useful for all learners and not only for aspiring scientists. Try to link the practical work as much as possible with their daily life and preconceptions.

3. Organization, analysis, and interpretation of data

Once collected, data must be ordered in a form that can reveal patterns and relationships and allows results to be communicated to others. We list goals about analysing and interpreting data. By the end of secondary education, students should be able to:

- Analyze data systematically, either look for relevant patterns or test whether data are consistent with the initial hypothesis.
- Recognize when data conflict with expectations and consider what revisions in the initial model are needed.
- Use spreadsheets, databases, tables, charts, graphs, statistics, mathematics, and ICT to compare, analyze, summarize, and display data and explore relationships between variables, especially those representing input and output.
- Evaluate the strength of a conclusion that can be inferred from any data set, using appropriate grade-level mathematical and statistical techniques.
- Recognize patterns in data that suggest relationships worth investigating further. Distinguish between causal and correlational relationships.
- Collect data from physical models and analyze the performance of a design under a range of conditions.

4. Organizing lab experiments

- **Methods to organize practical work**

There are 3 methods of organizing practical work:

a. Each group does the same experiments at the same time

All learners can follow the logical sequence of the experiments, but this implies that a lot of material is needed. The best group size is 3, as all learners will be involved. With bigger groups, you can ask to do the experiment twice, where learners change roles.

b. Experiments are divided among groups with group rotation

Each group does the assigned experiment and moves to the next experiment upon a signal by the teacher. At the end of the lesson, each group has done every experiment. This method saves materials but is not perfect when experiments are not ordered in a logical way. In some cases, the conclusion of an experiment provides the research question for the next experiment. In that case, this method is not very suitable.

The organization is also more complex. Before starting the lesson, the materials for each experiment should be placed in the different places where the groups will work. Also, the required time for each experiment should be about the same. Use a timer to show learners the time left for each experiment. Provide an extra exercise for fast groups.

c. All experiments are divided among groups without group rotation

Each group does only one or two experiments. The other experiments are done by other groups. Afterward, the results are brought together and discussed with the whole class. This saves time and materials, but it means that each learner does only one experiment and 'listens' to the other experiments' description. The method is suitable for experiments that are optional or like each other. It is not a good method for experiments that all learners need to master.

- **Preparation of a practical work**
  - Have a look at the available material at school and make a list of what you can use and what you need to improvise.
  - Determine the required quantities by determining the method (see above).

- Collect all materials for the experiments in one place. If learners' group is small, they can come to get the materials on that spot, but with more than 15 learners, this will create disorder. In that case, prepare for each group a set of materials and place it on their desk.
- Test all experiments and measure the required time for each experiment.
- Prepare a nice but educational extra task for learners who are ready before the end of the lesson.
- Write on the blackboard how groups of learners are formed.

- **Preparation of a lesson for practical work**

In the lesson plan of a lesson with practical work, there should be the following phases:

1. The introduction of the practical work or the 'excite' phase consists of formulation of a key question, discrepant event, or a small conversation to motivate learners and make connections with daily life and learners' prior knowledge.
2. The discussion of safety rules for the practical work:
   - Only work at the assigned place; do not walk around in the class if this is not asked.
   - Long hairs should be tied together, and safety eyeglasses should be worn for chemical experiments.
   - Only the material needed for the experiment should be on the table.
   - The practical work instructions: how groups are formed, where they get the materials, special treatment of materials (if relevant), what they must write down…
   - When the practical work materials aren't yet at the correct location, then distribute them now. Once learners have the materials, it is more difficult to get their attention.

- **How to conduct a practical work**
   - Learners do the experiments, while the teacher coaches by asking questions (Explore phase).

- The practical work should preferably be processed immediately with an explain phase. If not, this should happen in the next lesson.

- **How to conclude the lesson of a practical work:**
  - Learners refer to instructions and conduct the experiment,
  - Learners record and interpret recorded data,
  - Cleaning the workspace after the practical work (by the learners as much as possible).

- **Role and responsibilities of teacher and learners in lab experiment**

The roles and responsibilities of teacher during a lab experiment

Before conducting an experiment, the teacher will do the following:

- Decide how to incorporate experiments into class content best,
- Prepare in advance materials needed in the experiment,
- Prepare protocol for the experiment,
- Perform in advance the experiment to ensure that everything works as expected,
- Designate an appropriate amount of time for the experiment - some experiments might be adapted to take more than one class period, while others may be adapted to take only a few minutes.
- Match the experiment to the class level, course atmosphere, and your students' personalities and learning styles.
- Verify lab equipment before lab practices.
- Provide the working sheet and give instructions to learners during lab session.

During practical work, the teacher's role is to coach instead of helping with advice or questions. It is better to answer a learner's question with another question than to immediately give the answer or advice. The additional question should help learners to find the answer themselves.

- Prepare some pre-lab questions for each practical work, no matter what the type is.

- Try and start the practical work: start with a discrepant event or questions that help define the problem or questions that link the practical work with students' daily life or their initial conceptions about the topic.
- Use coaching questions during the practical work: 'Why do you do this?', 'What is a control tube?', "What is the purpose of the experiment?', 'How do you call this product?', 'What are your results?' etc.
- Use some questions to end the practical work: 'What was the meaning of the experiment?', 'What did we learn?', 'What do we know now that we didn't know at the start?', 'What surprised you?'etc.
- Announce the end of the practical work 10 minutes before giving learners enough time to finish their work and clean their space.

The Role of a lab technician during a laboratory-based lesson

In schools having laboratory technicians, they assist the science teachers in the following tasks:

- Maintaining, calibrating, cleaning, and testing the sterility of the equipment,
- Collecting, preparing and/or testing samples,
- Demonstrating procedures.

The learners' responsibilities in the lab work

During the lab experiment, learners have different activities to do; general learner's activities are:

- Experiment and obtain data themselves,
- Record data using the equipment provided by the teacher,
- Analyze the data often this involves graphing it to produce the related graph,
- Interpret the obtained results and deduct the theory behind the concept under the experimentation,
- Discuss the error in the experiment and suggest improvements,
- Cleaning and arranging material after a lab experiment.

5. Safety rules and precautions during lab experiments

Regardless of the type of lab you are in, there are general rules enforced as safety precautions. Each lab member must learn and adhere to the rules and guidelines set, to minimize the risks of harm that may happen to them within the working environment. These encompass dress' code, use of personal protection equipment, and general behaviour in the lab. It is important to know that some laboratories contain certain inherent dangers and hazards. Therefore, when working in a laboratory, you must learn how to work safely with these hazards to prevent injury to yourself and other lab mates around you. You must make a constant effort to think about the potential hazards associated with what you are doing and think about how to work safely to prevent or minimize these hazards as much as possible. Before doing any scientific experiment, you should make sure that you know where the fire extinguishers are in your laboratory, and there should also be a bucket of sand to extinguish fires. You must ensure that you are appropriately dressed whenever you are near chemicals or performing experiments. Please make sure you are familiar with the safety precautions, hazard warnings, and procedures of the experiment you perform on a given day before you start any work. Experiments should not be performed without an instructor in attendance and must not be left unattended while in progress.

A. Hygiene plan

A laboratory is a shared workspace, and everyone has the responsibility to ensure that it is organized, clean, well-maintained, and free of contamination that might interfere with the lab members' work or safety.

For waste disposal, all chemicals and used materials must be discarded in designated containers. Keep the container closed when not in use. When in doubt, check with your instructor.

B. Hazard warning symbols

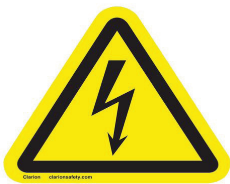
To maintain a safe workplace and avoid accidents, lab safety symbols and signs need to be posted throughout the workplace.

Chemicals pose health and safety hazards to personnel due to innate chemical, physical, and toxicological properties. Chemicals can be grouped into several different hazard classes. The hazard class will determine how similar materials should be stored and handled and what special equipment and procedures are needed to use them safely.

Each of these hazards has a different set of safety precautions associated with them.

The following table shows hazard symbols found in laboratories and the corresponding explanations.

Symbol		Meaning
	Highly flammable	Easily catches fire and burns
	Toxic	Can lead to death
	Irritant	Irritates the skin when in contact
	Radioactive	Dangerous to human health and can cause cancer
	Corrosive	Can easily burn you when in contact with your skin. It can also damage wood and metals

 Explosive	Can easily explode and releases small particles which can injure you
 Harmful	Poisonous when inhaled and ingested and can lead to death
 Electric chock	Can cause electric chock leading to death
 Laser radiations	Can cause cancer

C. General Laboratory Safety Rules.

You are ultimately responsible for your own safety and that of your fellow students, workers and visitors.

A standard list of basic laboratory safety rules are given below, and must be followed in every laboratory that uses hazardous materials or processes. These basic rules provide behavior, hygiene, and safety information to avoid accidents in the laboratory. Laboratory specific safety rules may be required for specific processes, equipment, and materials, which should be addressed by laboratory specific standard operating procedure.

1. The following Personal Protective Equipment must be worn at all times in the laboratory:

    - Lab coat.

- Eye protection: Chemical goggles. If you do get a chemical in your eye, rinse your eye immediately using large quantities of water or an eye wash bottle if available.
- Closed shoes with socks must be worn at ALL times – open-toed shoes, backless shoes and sandals are not permitted.
- Always wear gloves when working with unknown substances.
- Always wear the appropriate breathing masks when working with toxic or irritating vapours.

2. DO NOT work alone in a laboratory. Know the location and proper use of fire extinguishers, fire blankets and first aid kits.
3. Perform work with hazardous chemicals in a properly working fume hood to reduce potential exposures.
4. Always work in a well-ventilated area.
5. Working areas should be kept clean and tidy at all times.
6. Eating, smoking, and drinking are not allowed in a chemistry laboratory.
7. Labels and equipment instructions must be read carefully before use.
8. Long hair and loose clothing must be pulled back and secured from potential capture.
9. Avoid wearing jewellery in the lab as this can pose multiple safety hazards.
10. All containers must have appropriate labels. Unlabelled chemicals should never be used.
11. Do not taste or intentionally sniff chemicals.
12. Unused chemicals should not be returned to their original container unless directed to do so by the lab instructor.
13. DO NOT perform unauthorized experiments.
14. Never leave containers of chemicals open.
15. Avoid distracting or startling persons working in the laboratory.
16. Securely replace lids, caps, and stoppers after removing reagents from containers.
17. All flammable reagents must be removed before lighting a burner.

18. Never pour water into concentrated acid.
19. Mouth suction is never used to fill a pipette.
20. Always wipe spatulas clean before and after inserting into reagent bottles.
21. Report any accident and/or injury, however minor, to your instructor immediately.
22. Clean up any chemical spilled on the floor or any other working place immediately.
23. Before leaving the laboratory, make sure your work area is clean and dry and also ensure that all gas and water are completely turned off.
24. Wash exposed areas of the skin prior before leaving the laboratory.
25. Return materials used in the laboratory storage facility.
26. Never hesitate to ask questions especially if there is any question concerning proper operating procedure. Be sure that you understand every instruction before proceeding.
27. Never store food or beverages or apply cosmetics in areas where chemicals are used

6. Guidance on the Management of lab materials: Storage Management, Repairing and Disposal of Lab equipment

- **Keeping and cleaning up**

Working spaces must always be kept neat and cleaned up before leaving. Equipment must be returned to its proper place. Keep backpacks or bags off the floor as they represent a tripping hazard. Open flames of any kind are prohibited in the laboratory unless specific permission is granted to use them during an experiment.

- **Management of lab materials**

A science laboratory is a place where basic experimental skills are learned only by performing a set of prescribed experiments. Safety procedures usually involve chemical hygiene plans and waste disposal procedures. When providing chemicals, you must read the label carefully before starting the experiment. To avoid contamination and possibly violent reaction, do never return unwanted chemicals to their container. In the laboratory, chemicals should be stored in their original containers, and cabinets should be suitably ventilated. It is important to notify students that chemicals cannot be stored in containers on the floor. Sharp and pointed tools should be stored properly.

Students should always behave maturely and responsibly in the laboratory or wherever chemicals are stored or handled.

– **Hot equipment and glassware handling**

Hazard symbols should be used as a guide for the handling of chemical reagents. Chemicals should be labelled as explosives, flammable, oxidizers, toxic and infectious substances, radioactive materials, corrosives etc. All glassware should be inspected before use, and any broken, cracked, or chipped glassware should be disposed of in an appropriate container. All hot equipment should be allowed to cool before storing it.

All glassware must be handled carefully and stored in its appropriate place after use. All chemical glass containers should be transported in rubber or polyethylene bottle carriers when leaving one lab area to enter another. When working in a lab, do never leave a hot plate unattended while it is turned on. It is recommended to handle hot equipment with safety gloves and other appropriate aids but never with bare hands. You must ensure that hands, hair, and clothing are kept away from the flame or heating area and turn heating devices off when they are not in use in the laboratories.

– **Waste disposal considerations**

Waste disposal is a normal part of any science laboratory. As teachers or students perform demonstrations or laboratory experiments, chemical waste is generated.

These wastes should be collected in appropriate containers and disposed of according to local, state, and federal regulations. All schools should have a person with the responsibility of being familiar with this waste disposal. In order to minimize the amount of waste generated and handle it safely, there are several steps to consider.

Sinks with water taps for washing purposes and liquid waste disposal are usually provided on the working table. It is essential to clean the sink regularly. Notice that you should never put broken glass or ceramics in a regular waste container. Use a dustpan, a brush, and heavy gloves to carefully pick-up broken pieces, and dispose of them in a container specifically provided for this purpose. Hazardous chemical waste, including solvents, acids, and reagents, should never be disposed of down sewer drains. All chemical waste must be identified properly before it can be disposed of. Bottles containing chemical waste must be labeled appropriately. Labelling should include the words "hazardous waste." Chemical waste should be disposed of in glass or polyethylene bottles. Plastic coated glass bottles are best for this purpose. Aluminium cans that are easily corroded should not be used for waste disposal and storage.

- **Equipment Maintenance**

Maintenance consists of preventative care and corrective repair. Both approaches should be used to keep equipment in working order. Records of all maintenance, service, repairs, and histories of any damage, malfunction, or equipment modification must be maintained in the equipment logs. The record must describe hardware and software changes and/or updates and show the dates when these occurred. Each laboratory must maintain a chemical inventory that should be updated at least once a year.

7. Student experiment work sheet

There should be a sheet to guide students about how they will conduct the experiment, materials to be used, procedures to be followed and the way of recording data. The following is a structure of the student experiment worksheet. It can be prepared by teacher or be availed from the other level.

1. Date
2. Name of student/group
3. The title of experiment
4. Type of experiment (concept, equipment and inquiry based)
5. Objective(s) of the experiment
6. Key question(s)
7. Materials (apparatuses and chemicals, resources, etc…)
8. Procedures & Steps of experiment
9. Data recording and presentation

Test	Results/ observations	Comments
1		
2		
3		
Etc; ….		

10. Reflective questions and answers

    Question1

    Question 2

    Question 3

11. Answer for the key question.

8. Report Template for Learner

After conducting a laboratory experiment, students should write a report about their findings and the conclusion they took.

The report to be made depends on the level of students.

The following is a structure of the report to be made by a group of secondary school learners (S1-S3).

1. Introduction (details related to the experiment: Students identification, date, year, topic area, unit title and lesson).
2. The title of experiment.
3. Type of experiment (concept, equipment and inquiry based)
4. Objective(s) of the experiment.
5. Key question(s)
6. Materials (apparatuses and chemicals, resources, etc…)
7. Procedures & Steps of experiment
8. Data recording
9. Data analysis and presentation (Plots, tables, pictures, graphs)
10. Interpretation/discussion of the results, student alternative ideas form observation.

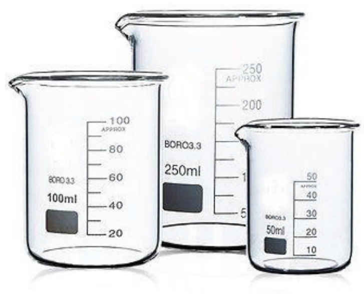
11. Theory or Main ideas concept, formulas, and application.

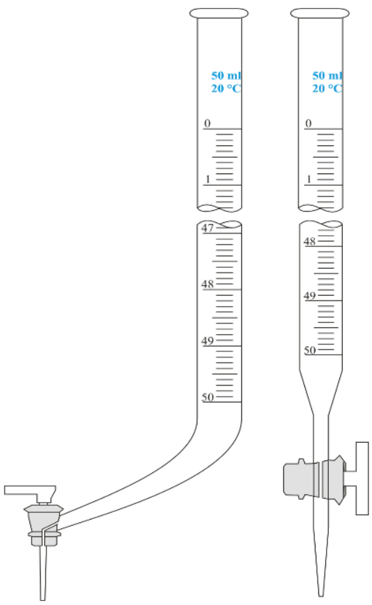
12. Conclusion (answer reflective questions and the key question).

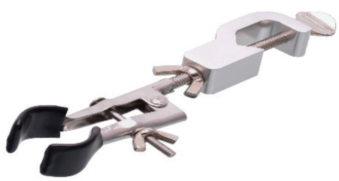
As a conclusion, there are safety rules and precautions to consider before, during and at the end of a lab experiment. We hope teachers are inspired to conduct lab experiments in a conducive Competence Based Curriculum way.

PART 2: LIST OF MAIN KIT ITEMS AND LAB MATERIALS NEEDED IN SCHOOLS

A. List of Apparatus

#	Item and description	Picture	Description of uses
1	Beaker		Used to hold and heat liquids. Multipurpose and essential in the laboratory.
2	Brushes		Used to easily clean the inside of a test tubes and other glassware.
3	Buchner funnel		Used with vacuum flask for performing vacuum filtration.

4	Bunsen burner		First, make sure your workspace is free of potential fire hazards. Once you are sure space is safe, connect the gas line and ignite the burner. Adjust the metal collar and needle gas valve at the burner's base to control the flame's size and temperature. When you're finished, close the air and gas ports, shut off the gas main line, and put the burner away once it's cool. Bunsen burner is used for heating and exposing items to flame.
5	Burette		Operating the Burette Proper burette technique is an important laboratory skill that may take some practice to develop. Although it may seem initially awkward, a right handed person should operate the burette with the left hand, and a left handed person should operate the burette with the right. This leaves your more coordinated hand to swirl the reaction flask if needed.

			Before delivering any solution, record the initial burette reading in your notebook. Open the stopcock by twisting it 90 degrees into the vertical position and allow the solution to drain. As you are near the desired volume, slow the flow by turning the stopcock back toward the closed position. You should be able to control the burette to deliver one drop at a time. When the desired volume has been delivered, close the stopcock.Wait a couple of seconds, then record the final burette reading.
6	Burette clamps	 	Used to hold burettes on a ring stand.

7	Clay triangle		Used to hold crucibles when they are being heated. They usually sit on a ring stand.
8	Crucible with lid		Used to heat small quantities to very high temperatures.
9	Crucible tong		Used to hold crucibles and evaporating dishes when they are hot.
10	Disposable pipette		Used for moving small amounts of liquid from place to place. They are usually made of plastic and are disposable.
11	Electronic balance		Used for weighing substances or objects, usually in grams. Place the electronic balance on a flat, stable surface indoors.

			Press the "ON" button and wait for the balance to show zeroes on the digital screen. Place the empty container you will use for the substance to be measured on the balance platform. Press the "Tare" or "Zero" button to cancel automatically the weight of the container. The digital display will show zero again. Carefully add the substance to the container. Ideally this is done with the container still on the platform, but it may be removed if necessary. Avoid placing the container on surfaces that may have substances which will add mass to the container such as powders or grease. Place the container with the substance back on the balance platform if necessary and record the mass as indicated by the digital display.

12	Erlenmeyer flask/Conical flask		Used to heat, mix, and store liquids. The advantage to the Erlenmeyer flask is that the bottom is wider than the top so it will heat quicker because of the greater surface area exposed to the heat.
13	Evaporating dish		Used to recover dissolved solids by evaporation.
14	Forceps		Used for picking up and moving small objects.
15	Glass funnel & Polypropylene funnel	 	Used to pour liquids into any container so that they will not be lost or spilled. They are also used with folded filter paper for filtration.

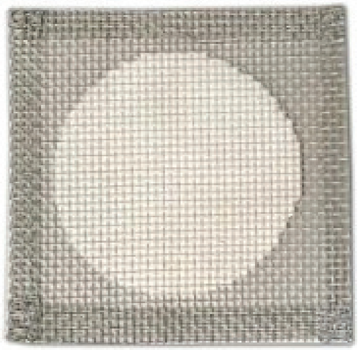
16	Glass stir rod		Used to stir liquids. They are usually made of glass.
17	Graduated cylinder/ measuring cylinder		Used to measure the volumes of liquids.
18	Micropipette		Used for accurately measuring and delivering very small volumes of liquid-usually 1 mL or less. Steps to follow when using a micropipette Select the volume. Set the tip. Press and hold the plunger at the first stop. Place the tip in the liquid. Slowly release the plunger. Pause for a second and then move the tip.

			Insert the tip into the delivery vessel. Press the plunger to the second stop.
19	Mortar and pestle		Used to crush solids into powders for experiments, usually to better dissolve the solids.
20	Pipette filler		How does a pipette filler work? Siphon liquid into the pipette to the desired level by squeezing valve "S" on the bottom of the pipette filler. This uses the vacuum created in the bulb to draw liquid into the pipette. Be careful not to draw liquid into the pipette filler. ... This allows you to release liquid at the desired rate and to the desired level.
21	Pipette with pump		Used for accurately measuring and delivering small volumes of liquid- usually 0.1-10 mL.

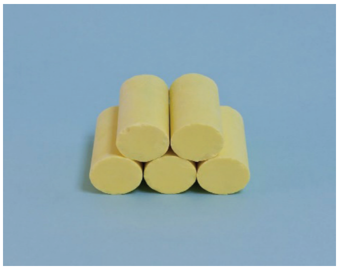
22	Ring clamp		Attached to a retort stand and with wire gauze used to hold beakers or flasks while they are heated by a gas burner.
23	Retort stand		Used to hold items being heated. Clamps or rings can be used so that items may be placed above the lab table for heating by Bunsen burners or other items. Used also to hold burette
24	Rubber stopper		Stoppers come in many different sizes. The sizes are from 0 to 8. Stoppers can have holes for thermometers and for other probes that may be used.
25	Separator funnel		For separating layers of immiscible liquids or for dropping liquids.

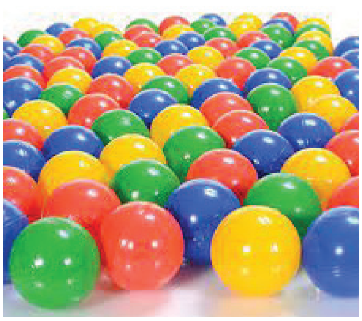
26	Spatula		Used for moving small amounts of solid from place to place.
27	Test tube		Used for storing, mixing, and heating small amounts of chemicals.
28	Test tube holder		Used to hold test tubes while heating.
29	Test tube rack		Used to hold test tubes while reactions happen in them or while they are not needed.

30	Thermometer		Used to take temperature of solids, liquids, and gases.
31	Vacuum filter flask		Used with vacuum line and Buchner funnel for vacuum filtration.
32	Volumetric flask		Used to prepare solutions with accurate concentration.
33	Wash bottle		Used to wash; rinse containers

34	Watch glass		Used to hold solids when being weighed or transported. They should never be heated. Can also be used to cover beakers or other containers.
35	Wire gauze		Used with a ring clamp to support glassware over a Bunsen burner.
36	Borosilicate glass tube		Used to connect to other items of glassware or equipment to deliver chemicals, solvents, liquids, gases and other products.
37	Deflagrating spoon or gas jar spoon		Generally used for the burning of materials inside gas jars or similar.

38	Thistle funnel		Thistle tubes are typically used to add liquid to an existing system or apparatus. Thistle funnels are used to add small volumes of liquids to an exact position.
39	Cardboard cover		Used to cover apparatus containing liquids
40	Rubber tube		Rubber tubing is often connected to a condenser, which is a laboratory tool used in the process of distillation. The rubber tubing helps cool water to flow in and out of the condenser and helps the heated water vapor in the condenser return to its liquid state.
41	Borosilicate delivery tube		The delivery tube is particularly useful for bubbling a gas from a gas cylinder or stoppered vessel through a liquid.

42	Trough		Used for collecting gases, such as hydrogen, oxygen and nitrogen. Troughs require a liquid such as water.
43	Beehive shelf		Usually used to support a receiving jar or tube while a gas is being collected over water with a pneumatic trough.
44	Graphite rods		Used as electrodes
45	Metallic rod		Used to test the heat conduction of metals
46	Aluminium foil		Can be used for temporary covering of instruments, shielding in vacuum equipment, packaging, wrapping, weighing boats, etc.
47	Sulphur rod		Used to test the heat conduction of non-metals

48	Charcoal rod		Used as electrodes
49	Syringe		Often used for measuring and transferring solvents and reagents where a high precision is not required.
50	Electrolyser		Used in the electrolysis process
51	Gas jar and cover		Used for collecting gas from experiments.
52	Plastic balls		Used to illustrate Brownian movement

B. List of chemicals

SN	CHEMICALS
1	Aluminium
2	Aluminium foil
3	Aluminium oxide, $Al_2O_3(s)$
4	Ammonia
5	Ammonium chloride
6	Ammonium nitrate
7	Anhydrous copper (II) sulphate
8	Silver nitrate, ($AgNO_3$)
9	Barium chloride
10	Barium nitrate
11	Barium sulphate
12	Black ink
13	Blue litmus papers
14	Bromine water
15	Calcium
16	Calcium carbonate
17	Calcium chloride (CaCl2)
18	Calcium hydroxide
19	Calcium oxide
20	Calcium sulphate
21	Candle wax
22	Charcoal rod
23	Chloroform
24	Cooking oil
25	Copper
26	Copper (II) carbonate
27	Copper (II) hydroxide

28	Copper (II) oxide
29	Copper (II) sulfate crystals
30	Copper electrode
31	Copper powder
32	Copper turnings
33	Sulphuric acid
34	Ethanoic acid
35	Ethanol
36	Graphite rods
37	Hydrochloric acid
38	Hydrogen peroxide
39	Iodine
40	Iron fillings
41	Iron powder
42	Iron rod
43	Lead (II) bromide
44	Lead (II) carbonate
45	Lead (II) chloride
46	Lead (II) nitrate
47	Limewater
48	Magnesium oxide
49	Magnesium ribbon
50	Magnesium sulphate
51	Manganese dioxide
52	Methyl orange
53	Naphthalene
54	Nitric acid
55	Oil
56	Phenolphthalein

57	Potassium
58	Potassium bromide
59	Potassium chloride
60	Potassium dichromate (VI)
61	Potassium hydroxide
62	Potassium iodide
63	Potassium nitrate
64	Potassium permanganate
65	Propanone
66	Red litmus paper
67	Silver bromide
68	Silver chloride
69	Silver nitrate
70	Sodium (Na)
71	Sodium acetate
72	Sodium carbonate
73	Sodium chloride
74	Sodium hydrogen carbonate
75	Sodium hydroxide
76	Sodium metal
77	Sodium sulphate
78	Sodium sulphite, Na_2SO_3
79	Sodium thiosulphate $Na_2S_2O_3$
80	Sugar
81	Sulphur powder
82	Sulphur rod
83	Sulphuric acid
84	Universal indicator
85	Wood charcoal

86	Zinc (Zn)
87	Zinc carbonate
88	Zinc granules
89	Zinc metal
90	Zinc powder
91	Cobalt (II) chloride paper

PART 3: DETAILED EXPERIMENTS PER GRADE OF O'LEVEL

CHEMISTRY EXPERIMENTS FOR SENIOR ONE

UNIT 2 Laboratory safety and apparatus

Experiment 2.1: Use of a fire extinguisher in case of fire in laboratory

Rationale: Fire extinguishing using a fire extinguisher is a first aid response to fire outbreak and can help prevent catastrophic damages to properties (laboratory, offices, classrooms,…) and even loss of life.

Objective

Learners will be able to use a fire extinguisher correctly.

Required materials

- Fire extinguisher

http://www.mscfire.ie/product/dry-powder-fire-extinguishers/

- Match box with sticks

https://www.indiamart.com/proddetail/match-box-only-for-odisha-15129377873.html

- Tiny firewood or trash papers/rubbish/grasses

Experiment set up

Source: *https://theconversation.com/how-does-the-stuff-in-a-fire-extinguisher-stop-a-fire-120859*

Procedures

1. Make fire using small pieces of firewood or trash papers/rubbish/grasses
2. Take the fire extinguisher
3. Pull the pin
4. Approach the fire standing at a distance
5. Point the nozzle or hose of the extinguisher at the base of the fire
6. Hold down the trigger
7. Sweep from side to side at the base of the fire, the fuel source, until the fire is out

Guiding questions

i. Why do you have to pull the pin out from the lever of a fire extinguisher every time you want to use it?

   Answer: The gas to be evolved and extinguishes a fire.

ii. What do you think would happen if the hose/pipe of the fire extinguisher has some cuts?

   Answer: *the gas will be released and should cause burning.*

iii. What to avoid while using fire extinguishers?

Interpretation

Fire extinguishers are portable or movable apparatus used to put out a small fire by directing onto it a substance that cools the burning material, deprives the flame of oxygen or interferes with the chemical reactions occurring in the flame.

Experiment 2.2: Measuring volume by using a measuring cylinder

Rationale: when working in chemistry labs, accuracy in measurements of reactant quantities is the need. This accuracy is also needed in different measurements in real life in factories, restaurants, bars even at home, using efficiently given containers.

Objective

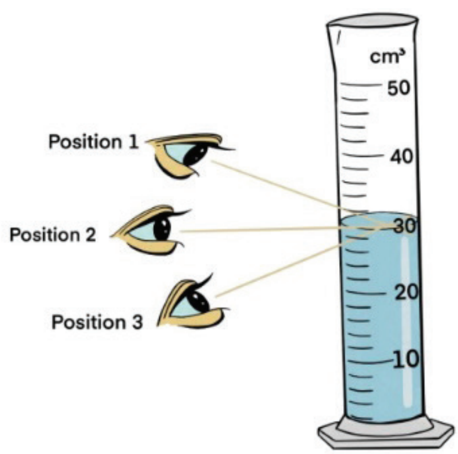
Learners will be able to use efficiently a measuring cylinder to measure volumes of liquids.

Required materials

Apparatus
- Measuring Cylinder
- Beaker

Chemical
- Water

Experiment set up



Procedure

1. Select a clean and dry measuring cylinder that is large enough to hold a required volume of liquid being measured;
2. Pour the liquid you want to measure into the measuring cylinder with care;
3. Hold the cylinder at eye level to take a reading;
4. Take the liquid measurement at the very bottom of the meniscus on the surface of the liquid (position 2 in the experiment set up above);
5. Read the measurement to which the graduation line of the meniscus is closest.

Evaluation activity

Use different measuring cylinders to measure the following volumes of water ($30cm^3$, $85\ cm^3$,$56\ cm^3$,$74\ cm^3$)

Data collection

Volume to be measured	Volume measured
$30Cm^3$	
$85\ Cm^3$	
$56\ Cm^3$	
$74\ Cm^3$	

Experiment 2.3: Separation of immiscible liquids by using a separating funnel

Rationale: In chemistry there are many methods to separate constituents from a given mixture. The method used will depend on the type of mixtures. To separate beans and stones, sand and water, banana beer and sorghum, …..appropriate methods are needed. According to the types of mixtures, a separating funnel will be used when constituents of the mixture are not miscible i.e water and oil, …

Objective: Learners will be able to separate immiscible liquids using a separating funnel.

Required materials

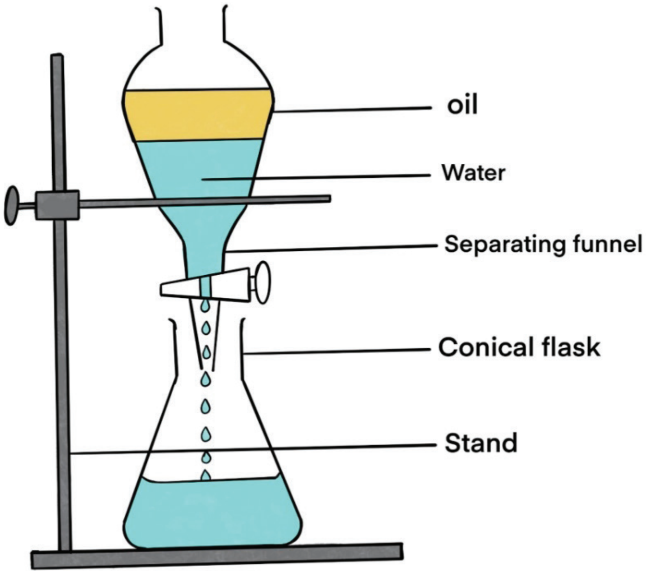
Apparatus

- Beakers
- Separating funnel
- Retort stand and clamp
- Conical flask

Chemicals

- Oil
- Water

Experiment set up



Procedure

1. Set up the apparatus as shown above,
2. Close the stopcock on the separating funnel,
3. Pour the mixture of oil and water into the separating funnel and allow the layers to separate,
4. Open gently the stopcock to let the bottom layer flow into the conical flask,

5. Close the stopcock and take away the conical flask containing water,
6. Place another conical flask under separating funnel,
7. Open gently the stopcock to let oil flow into the conical flask.
8. Why water is in the bottom layer?

   Answer: *Water is in the bottom layer because it is denser than oil.*

Interpretation of the results and Conclusion

Guiding questions

1. Are oil and water immiscible liquids?
2. What happens when immiscible liquids are mixed in a separating funnel?
3. Which liquid is collected firstly from the mixture of immiscible liquids?

Oil and water are immiscible liquids. In a separating funnel immiscible liquids form two layers. The denser liquid will be in the bottom layer and therefore collected firstly from the mixture. A separating funnel is used to separate immiscible liquids from a mixture.

Evaluation activity.

1. By using a separating funnel separate the components of a mixture of:

   a) Vinegar and oil,

   b) Kerosene and water.
2. Why cannot you use a separating funnel to separate the components of a mixture composed by ethanol and water.

   Answer: *Water and ethanol are miscible liquids*

Experiment 3: Measuring the temperature by using a thermometer

Rationale: Measuring accurately temperatures using thermometer is an important exercise. A number of diseases are characterised by a change of temperature in body.

Measuring body temperature is very important in medicine. In some illnesses, the evolution of the disease can be followed by measuring body temperature. This allows the doctor to analyse the effectiveness of treatments based on body temperatures.

Objective

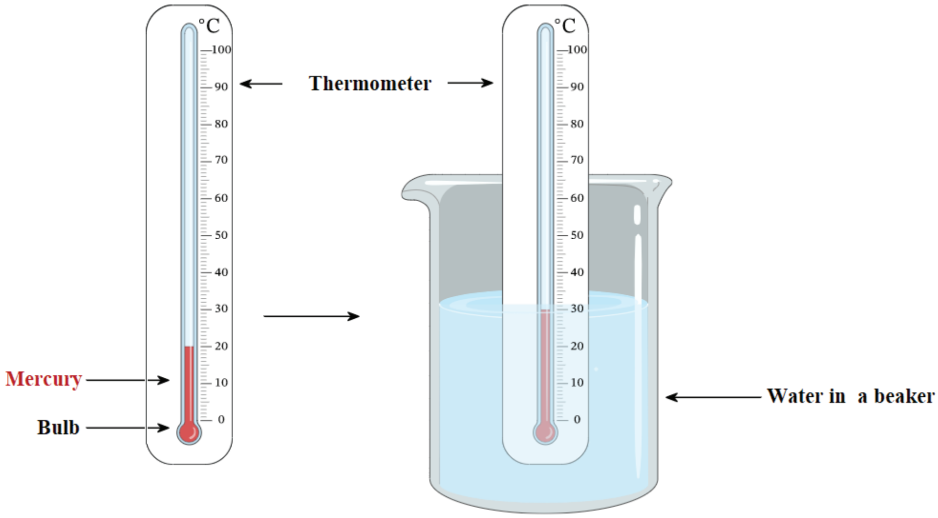
Learners will be to able measure efficiently the temperature of a substance using a thermometer.

Required materials

Apparatus
- Bunsen burner
- Beaker
- Thermometer
- Tripod
- Wire gauze
- Match box

Chemical
- Water

Experiment set up



Procedure

1. Pour cold water in a beaker,
2. Place in a beaker the bulb end of the thermometer for two minutes and record the temperature corresponding to the level of the mercury in the thermometer,
3. Record the temperature in Celsius degrees,
4. Heat the water in the beaker for two minutes while stirring with the same thermometer,
5. Record the new temperature of water,
6. Determine the temperature change of the water.

Evaluation activity

1. Take various liquids and heat them for three mutes to determine their temperature changes. Find the temperature of different liquids by using a thermometer and record the results.

Liquid	Initial Temperature	Final Temperature	Change
1. Ethanol			
2. Cyclohexane			
3. Juice			
4. Vinegar			

Experiment 2.4: Heating by using a Bunsen burner

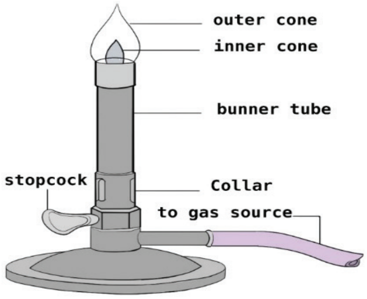
Rationale: Using Bunsen burner when heating requires many precautions such as working in a clean space, light the match before turning on the gas, turn off the gas after the use. These precautions are identical to those of heating using gas stove. Nowadays, Rwanda is encouraging people to use gas instead of charcoal or trees. This reduces pollution, time wastage while collecting firewood, save on the environment (reduces deforestation), reduces global warming. Having skills of using Bunsen burner implicate adequate use of gas stove.

Objective: Learners will be able to efficiently use a Bunsen burner.

Required materials

Apparatus	Chemical
– Bunsen burner	– Water
– Match box	

Experiment set up



Procedure

1. Place the Bunsen burner on the laboratory worktable,
2. Remove all papers, notebooks, combustible materials, and excess chemicals from the area,
3. Check the gas tube/cylinder if there is not any leakage of the gas,
4. Fix the gas tube to the gas source and check if it is closed,
5. Close the air hole using collar,
6. Light the match on the top and turn on the gas to have a yellow flame,
7. Open collar for a blue flame which is perfect for heating,
8. Close the collar and turn off the gas after the use.

Evaluation activity

Light the Bunsen burner following the procedure above.

Experiment 2.5: Weighing 10g of sodium chloride by using an electronic balance

Rationale: Mass is often measured by measuring an unmeasured object with standardized forms of measurements such as milligrams, grams and kilograms. In the shop to sell necessary quantities of sugar, salt, flour, …. requires accuracy mass measurement. Measuring height and weight accurately is important *when checking child's health.* Height and weight measurements are used to calculate your body mass index, or BMI, a measure of healthy versus unhealthy weight.

Objective

Learners will be able to use an electronic balance efficiently.

Required materials

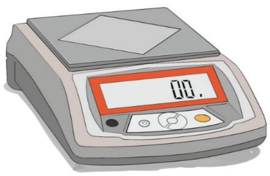
Apparatus

- Electronic balance
- Spatula
- Watch glass

Chemical

- Sodium chloride

Experiment set up



Procedure

1. Connect electronic balance to electric power source,
2. Turn on balance,
3. Place the watch glass on a balance and set to 0 by pressing the tare button,
4. Using a spatula, add 10g of sodium chloride to the watch glass.

Evaluation activity

Weigh 5g, 6g and 10g of calcium carbonate.

Unit 3: States and changes of states of matter

Experiment 3.1: Melting of ice

Rationale: Snow forms when the atmospheric temperature is at or below freezing (0 degrees Celsius or 32 degrees Fahrenheit) and there is a minimum amount of moisture in the air. If the ground temperature is at or below freezing, the snow will reach the ground. Why snow does not persist indefinitely on the ground. As temperatures increase above freezing point, heat from the sun the snow begins to melt. This natural phenomenon is clearly explained by the melting of ice process.

Objective

Learners will be able to show and explain the melting of substances.

Required materials

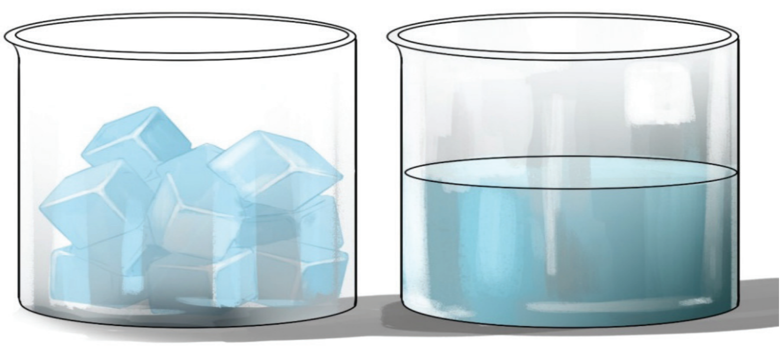
Apparatus

- Beaker
- Heater

Chemical

- Ice

Experiment set up



Ice *Liquid water*

Procedures

1. Take some pieces of ice,
2. Put them in a beaker,
3. Heat the beaker for three minutes while observing the change.

   What do you observe?

   Answer: ***Ice crystals in the beaker change into liquid water when heated.***

Interpretation of results and conclusion

Solid substances generally change into liquids when their temperatures are increased. This change of physical states is called **Melting.**

Experiment 3.2: Evaporation of water

Rationale: The water cycle is a natural phenomenon that includes many different processes: liquid water from oceans, lakes and rivers evaporates into water vapors, condenses to form clouds, and precipitates back to earth in the form of rain and snow. Evaporation of water is then the first step of the cycle.

Objective

Learners will be able to show and explain the evaporation of substances.

Required materials

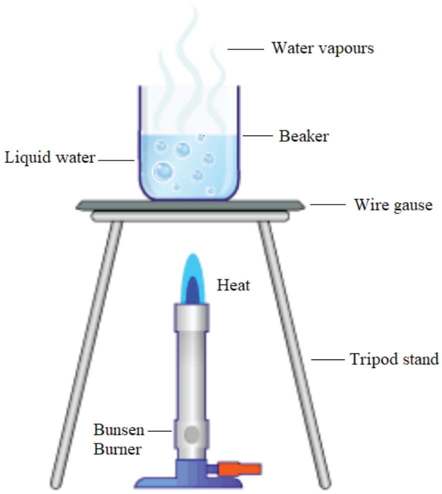
Apparatus

- Bunsen burner
- Borosilicate beaker
- Tripod stand
- Wire gauze
- Match box

Chemical

- Water

Experiment set up



Procedure

1. Pour water in a beaker,
2. Put the beaker that contains water on tripod stand,
3. Heat for four minutes.

   What do you observe?

   Answer: *Some water vapours are produced after some time.*

Interpretation of results and Conclusion

When liquid substances are heated, they change to vapours or gases. This change of state from liquid to gas is called **evaporation.**

Experiment 3.3: Sublimation and Deposition

Rationale: Depositional landforms is an example of deposition in geography. These landforms are the visible evidence of processes that have deposited sediments or rocks after they were transported by flowing ice or water, wind or gravity. Examples include beaches, deltas, glacial moraines, sand dunes and salt domes.

- **About sublimation:** naturally dry ice sublimes, snow and ice sublime during winter season without melting, room fresheners which are used in toilets sublimes. All these processes are showing that we can find sublimation and deposition process are natural.

Objective:

Learners will be able to show and explain the change of states from solid to gas and from gas to solid.

Required materials

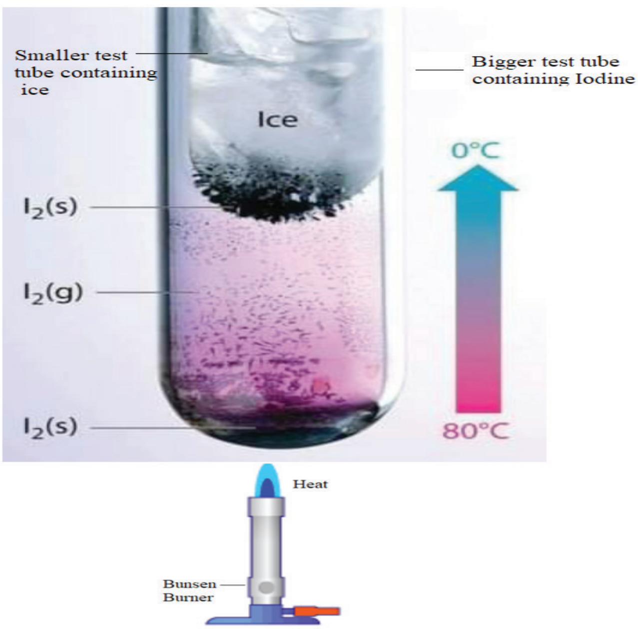
Apparatus

- Bunsen burner
- Boiling tubes
- Spatula
- Electronic balance
- Retort stand and clamp
- Match box

Chemical

- Iodine

Experiment set up



Procedure

1. Place three spatulas full of iodine in big boiling tube.
2. Fix the boiling tube on the retort stand.
3. Heat the boiling tube gently over a non-luminous flame of the Bunsen burner.
4. Insert a smaller boiling tube containing ice into the bigger one.

What do you observe?

Answers

i. When Iodine is heated, purple vapours are observed.

ii. The purple vapours solidify on a cool surface.

Interpretation of results and Conclusion

Some solid substances change directly to gases without passing through the liquid state. This change of state from solid to gas state is called **sublimation.** It is also possible for a gas to change directly into a solid and this change of state is called **deposition.**

Experiment 3.4: Condensation

Rationale: The water cycle is a natural phenomenon that includes many different processes. One of these processes is condensation of water vapors to form clouds and formation of rain and/or snow. The condensation explains also the formation of rain from water of oceans, lakes and rivers.

Objective

Learners will be able to demonstrate and explain the change of state from gas to liquid.

Required materials

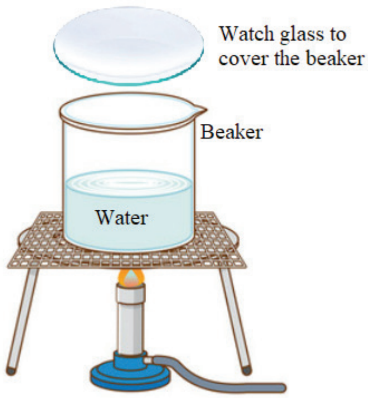
Apparatus

- Beaker
- Bunsen burner
- Tripod stand

Chemical

- Water

Experiment set up



Procedure

1. Pour water into a beaker and cover it with a watch glass.
2. Place it on a tripod stand.
3. Heat the water in the beaker for some minutes.

   What do you observe?

   Answer: Water vapours are produced inside the beaker and after a while, droplets of water are observed on the internal wall of the watch glass.

Interpretation of results and Conclusion

When vapours of substances are cooled, they are transformed into liquids. This change of state from gas to liquid is called **condensation**.

Experiment 3.5: Difference between chemical and physical changes

Rationale: Reactions in chemistry are all chemical changes characterised by the formation of a substance having different properties from the original substance. Reactions that occur in our body, in factories when making different articles, changes that occur when cooking, participate in making new substances. But changes in the water cycle are all physical changes.

Objective

Learners will be able to explain the difference between physical and chemical changes.

Required materials

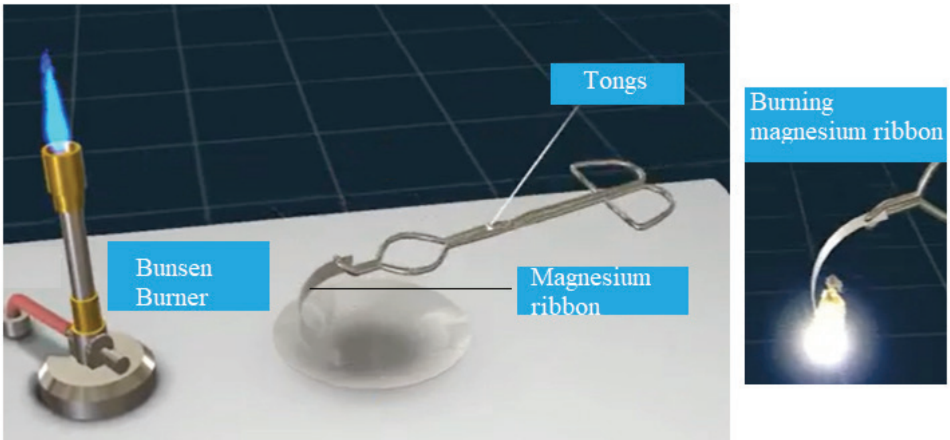
Apparatus

- Pair of tongs
- Bunsen burner
- Match box
- Watch glass

Chemicals

- Magnesium ribbon
- Water

Experiment set up



Caution:

Burning magnesium produces intense white light that can cause temporary loss of sight. Do not look directly at the flame of burning magnesium.

Procedure

1. Using a pair of tongs, heat a piece of magnesium ribbon on a Bunsen burner flame until it starts to burn.
2. Collect the product on the watch glass and compare it to the initial magnesium ribbon.

Questions

1. State the observations made when magnesium ribbon is burnt.
2. How is the product formed above different from the initial substance?
3. Is there new substance formed?
4. Are vapours or steams formed from liquid water different from water?

Answers

1. When magnesium is burnt in air, it produces a brilliant white flame and ash.
2. The magnesium ribbon turns into ash.
3. Yes, the ash of magnesium formed is a new substance.
4. Vapours of water or steams and water are the same substance with different physical states.

Interpretation of results and Conclusion

Guiding questions

1. What happens when metallic magnesium is burnt in air?
2. Is the powder formed different from magnesium ribbon? Explain your answer.
3. Name the compound formed when metallic magnesium is burnt in air.
4. Is the burning of metallic magnesium physical or chemical change? Explain.

When metallic magnesium is burnt in air, it forms a white solid which can be crushed into a powder. The powder formed is different from magnesium ribbon. The white powder formed is called magnesium oxide. It is not possible to get back the magnesium ribbon from the powder. This process is irreversible. The formation of magnesium oxide from magnesium is a **chemical change.** Evaporation process is a **physical change** as it is reversible.

Evaluation activity

Classify the following changes into chemical and physical changes: Ripening of fruits, deposition, freezing, fermentation of sugars, fusion, sublimation, rusting of iron, condensation, and sublimation.

Experiment 3.6: Brownian motion

Some natural examples of Brownian motion are movement of dust motes in a room (although largely affected by air currents), diffusion of pollutants in the air and diffusion of calcium through bones. Calcium is the major component of the bone, where it is present at more than 99% as calcium-phosphate complexes, and provides the skeleton strength and structure.

Objective

Learners will be able to demonstrate the random movement of particles suspended in a fluid.

Required materials

Apparatus

- Beaker
- Plastic balls
- Bunsen burner
- Match box

Chemical

- Water

Experiment set up

Procedure

1. Take a beaker and put some plastic balls in it.
2. Pour water in the beaker and boil the mixture on a Bunsen burner.
3. Observe the changes.

Questions

1. How is the motion of the plastic balls?

Answer

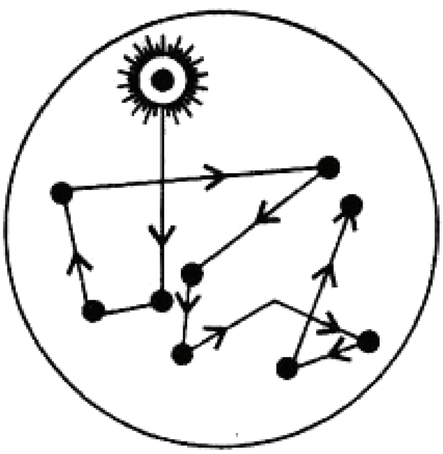
The plastic balls move randomly in different directions.

Interpretation of results and Conclusion

Guiding questions

1. What happened when water containing plastic balls is heated?
2. What is the cause of the random movement of these plastic movement?

When water containing plastic balls is heated, these plastic balls move randomly in different directions. This is also the case of vapour particles of water even though it cannot be easily seen. This random motion is called Brownian motion. Brownian motion is the random movement of particles suspended in a fluid. This is a result of the movement of the various particles which are constantly colliding which each other.



Other examples of Brownian motion include:

- The motion of pollen grains on still water,
- Movement of dust motes in a room (although largely affected by air currents),
- Diffusion of pollutants in the air.

Experiment 3.7: Diffusion of ammonia and hydrogen chloride

Rationale: The diffusion of gases is the process of movement of molecules from a region of higher concentration to a region of lower concentration. Diffusion is a very important process for photosynthesis where carbon dioxide from the stomata diffuses into the leaves and finally into the cells.

The diffusion of gases in plants through pores is sufficient to service the few living cells in its stem cortex and its thalloid leaves. In the roots, gas exchange is restricted to a small permeable area. Stomata are holes or empty gaps between guard cells.

Objective

Learners will be able to demonstrate and explain the diffusion process by using ammonia and hydrogen chloride.

Required materials

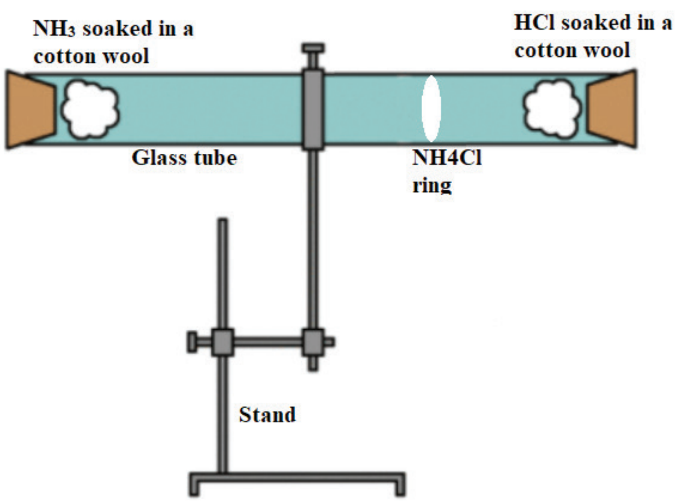
Apparatus

- Long glass tube
- Cotton wool
- Tongs
- Stand and clamp

Chemicals

- Concentrated ammonia
- Concentrated hydrochloric acid

Experiment set up



Caution

Ammonia and hydrogen chloride gases are poisonous. Point the bottle away from both you and the students.

Procedure

1. Set up the apparatus as shown above,
2. Soak two pieces of cotton wool, one piece in concentrated ammonia solution and the other in concentrated hydrochloric acid separately.
3. Quickly insert the soaked cotton wool pieces simultaneously at the opposite ends of the long glass tube.
4. Carefully observe what happens in the glass tube.
5. Measure the distance from both ends of the glass tube to the position where the ring of fumes is seen.

   What do you observe in the glass tube? Explain your observations.

   Answers

   1. Dense white fumes are formed inside the glass tube.
   2. Ammonia and hydrogen chloride gases diffuse in the long glass tube. When the two gases meet white fumes are formed. The white ring is formed closer to the end with cotton wool soaked in concentrated hydrochloric acid.

Interpretation of results and Conclusion

Guiding questions

1. What happens when the cotton wool is soaked in concentrated ammonia and when the cotton wool is soaked in concentrated hydrochloric acid?
2. What happens when the two gases, ammonia and concentrated hydrochloric acid meet?
3. Write a balanced equation of the reaction when ammonia and hydrochloric acid meet.
4. To which end the white ring formed is closer to?

The cotton wool soaked in concentrated ammonia solution gives out ammonia gas whereas the cotton wool soaked in concentrated hydrochloric acid gives out hydrogen chloride gas.

Ammonia and hydrogen chloride gases diffuse in the long glass tube. When the two gases meet, white fumes of ammonium chloride are formed. The ammonium chloride is seen as a white ring.

Equation of the reaction that occurs: HCl (g) + NH3(g) – NH4Cl(g): white fumes

The white ring is formed closer to the end with cotton wool soaked in concentrated hydrochloric acid. Ammonia gas has less dense particles. Therefore, its particles move faster than hydrogen chloride gas particles.

Experiment 3.8: Diffusion of potassium manganate (VII) in water

Objective

Learners will be able to demonstrate and explain the diffusion process of potassium manganate (VII)

Required materials

Apparatus

- Beaker

Chemicals

- Potassium manganate (VII)
- Water

Experiment set up

Procedure

1. Pour water into a beaker,
2. Add a half full spatula of potassium manganate (VII) in water.

   What do you observe?

   Answer: *The purple colour is slowly spreading into the water.*

Interpretation of results and conclusion

Guiding questions

1. What do you observe when you put some potassium manganate (VII) crystals in a beaker of water?
2. Explain why the potassium manganate (VII) spreads in water.
3. Explain from which region particles move generally.

When you put some potassium manganate (VII) crystals in a beaker of water, the potassium manganate (VII) particles spread until all the water is uniformly coloured. The potassium manganate (VII) spreads in water due to the movement of particles from potassium manganate (VII) into the water. Thus, generally particles move from the region where they are greater in number to the region where they are fewer in number by the process of diffusion.

Unit 4: Pure substances and mixtures

Experiment 4.1: Filtration

Rationale: Filtration is one of the methods used to separate constituents of a mixture in laboratories even in daily life. Among many other uses and other complementary methods, filtration is used in water treatment to give people access to clean water.

Required materials

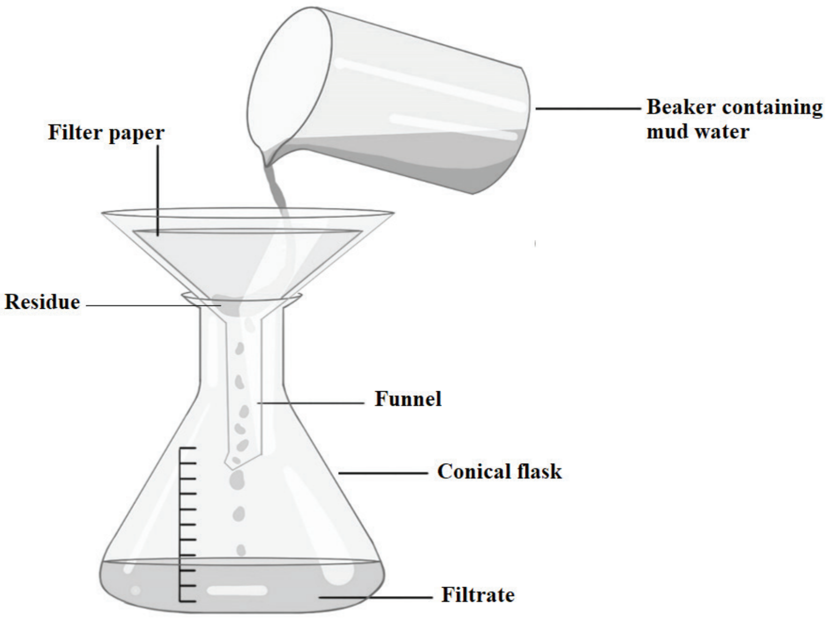
Apparatus

- Conical flasks
- Filter paper
- Funnel
- Beaker

Chemicals

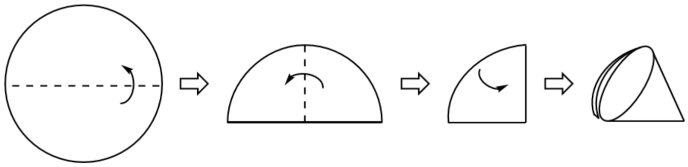
- Soil
- Water

Experiment set up



Procedure

1. Put the soil in a beaker containing water and stir using glass rod,
2. Fold a filter paper to form a quadrant, and then open it up into a hollow cone,



3. Wet the filter paper to make it stick on the funnel,
4. Put the funnel on the conical flask,
5. Stir the mixture of soil and water and pour it into the funnel fitted with the filter paper.

   Describe the contents left on the filter paper and those in the conical flask.

   Answer: *The content left on the filter paper is the one which cannot pass through the paper but the content in conical flask is the one that can pass through the filter paper.*

Interpretation of results and Conclusion.

Filtration is a method used to separate solid particles from a liquid. During the filtration, the liquid which passes through the filter paper is called the **filtrate** while the solid that remains on it is called the **residue**.

Evaluation activity

Give examples of mixtures whose constituents can be separated using filtration method.

Experiment 4.2: Distillation

Rationale

- Distillation is used in water treatment, contaminated water is heated to form steam. Inorganic compounds and large non-volatile organic molecules do not evaporate with the water and are left behind. The steam then cools and condenses to form purified water.
- Distilled water is used in refinery: Crude oil is made up of a mixture of hydrocarbons, and the distillation process separate this crude oil into its component hydrocarbons, or "fractions."
- Distillation is used to separate mixtures of liquids by exploiting differences in the boiling points of the different components. The technique is widely used in industry, for example in the manufacture and purification of nitrogen, oxygen and the rare gases.

Objective

Learners will be able to carry out and describe the distillation process.

Required materials

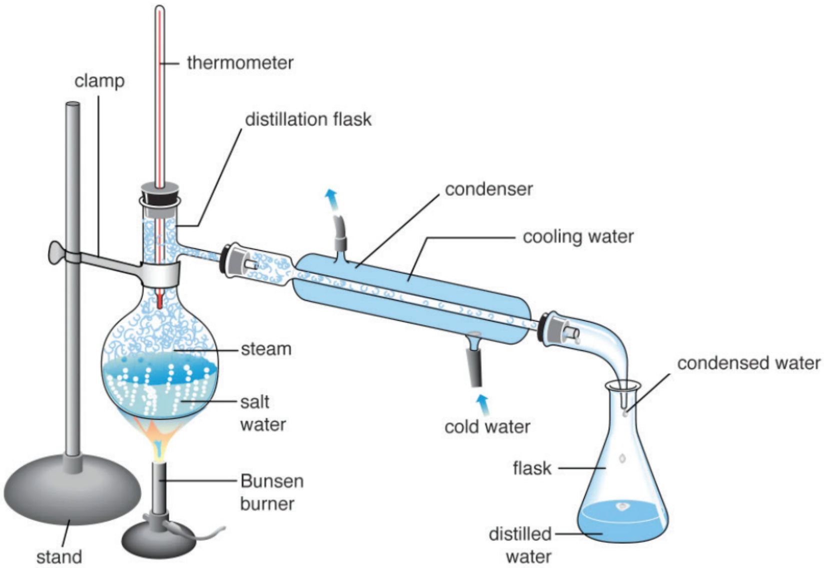
Apparatus

- Round bottomed flask
- Bunsen burner
- Retort Stand with accessories
- Condenser
- Wire gauze
- Thermometer
- Beaker
- Match box
- Rubber tubing

Chemicals

- Water
- Common salt

Experiment set up



Procedure

1. Dissolve 5g of table salt in 200 mL of water,
2. Pour the solution into a round bottomed flask,
3. Arrange the apparatus as shown in the set up above,
4. Heat up the solution in the round bottomed flask.

   What do you observe?

   Answer:

   The solution in the round bottom flask changes into vapours which condense into liquid in the condenser. The liquid is collected in the conical flask.

   When all the liquid is evaporated, the solid salt remains at the bottom of the flask.

Interpretation of results and conclusion

Guiding questions:

1. Describe the simple distillation process.
2. What is the role of the thermometer in this experiment?

The liquid collected in the conical flask is called **distillate**. This process is called **simple distillation.** During distillation, evaporation and condensation take place at the same time but in different pieces of apparatus.

Commonly a thermometer is used to note the temperature at which the solution boils if the solution is made by components which have boiling temperatures which are not highly different.

Experiment 4.3: Paper chromatography

Rationale

Paper chromatography is an analytical technique used to separate mixtures of chemicals (sometimes colored pigments) using a partitioning method. In laboratory some other mixtures which can be separated by this method are coloured dyes in flower extracts, components of blood, components of urine and dyes in food colours. But chromatography is also used to study the process of fermentation and ripening, to check the purity of pharmaceuticals, to inspect cosmetics and to detect the adulterants, to detect the contaminants in drinks and foods, to examine the reaction mixtures in biochemical laboratories: separation of mixtures of amino acids, peptides, carbohydrates, steroids and purines.

Objective

Learners will be able to use paper chromatography to separate the pigments found in ink samples.

Required materials

Apparatus

- Circular filter paper
- Beaker
- Dropper

Chemicals

- Propanone
- Black ink

Experiment set up

Procedure

1. Place a filter paper on the beaker,
2. Use a dropper to put a drop of black ink at the centre of the filter paper,
3. Add propanone drop gently using a different dropper over the black ink drop, waiting till each drop stops spreading.
4. Repeat until circular bands of colours appear,
5. Leave the paper to dry,
6. Record the colours that appear.

   Why do we use propanone?

   Answer: *The ink dissolves in propanone*

Interpretation of results and conclusion

Guiding questions

1. Explain the formation of several rings on the filter paper.
2. Why is propanone used in this experiment?

Ink is made of several dyes as evident from the formation of several rings formed. The different dyes in the ink dissolve in propanone and spread out through the filter paper to form a series of bands with various colours. The series of these bands of various colours corresponding to different components of the ink is called a **chromatogram**.

Some other mixtures which can be separated by this method are coloured dyes in flower extracts, components of blood, components of urine and dyes in food colours.

Experiment 4.4: Evaporation

Objective

Learners will be able to separate components of a mixture by using evaporation.

Required materials

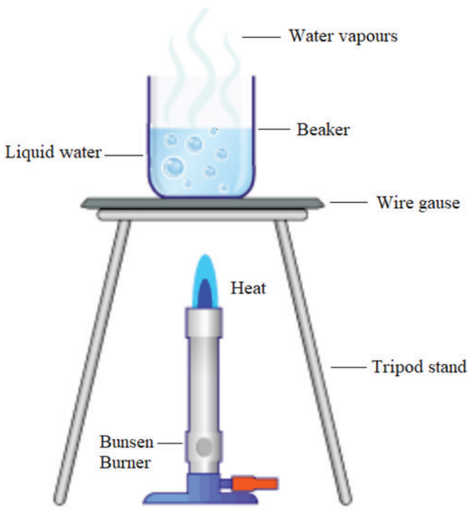
Apparatus

- 100mL borosilicate Beaker
- Bunsen burner
- Glass rod
- Evaporating dish

Chemicals

- Common salt
- Water

Experiment set up



Procedure

1. Put 10g of common salt in a 100mL borosilicate beaker half filled with water,
2. Stir with a glass rod to make a homogeneous solution,
3. Heat the mixture until all the solvent is evaporated,

   What do you observe in the beaker?

   Answer: After all the water has evaporated, white solid particles remain in the beaker.

Interpretation of results and conclusion

Guiding question:

1. Explain how evaporation process is used to separate the components of a homogeneous mixture.

When a solution of water and table salt is boiled, steam (water vapours) is produced, and it escapes from the solution. As time goes on the amount of water reduces gradually and eventually only solid crystals remain in the beaker.

The process of changing a liquid to gaseous form is called **evaporation**. This process is used to separate the components of a homogeneous mixture made up of a solid which has a high melting temperature and a liquid solvent of a low boiling temperature.

Experiment 4.5: Magnetic Separation of iron powder and sand.

Rationale of the experiment

Some mixtures can be separated using physical methods alone that is this experiment is will help us to know how to separate mixtures using magnet. Magnetic bar can be used to separate the mixture of iron powder and sand.

Objective

Learners will be able to separate a mixture of iron powder and sand correctly by using a magnet.

Required materials

Apparatus

- Magnetic bar
- Petri dish/ plastic plate/bowl Plastic bag/wrapper

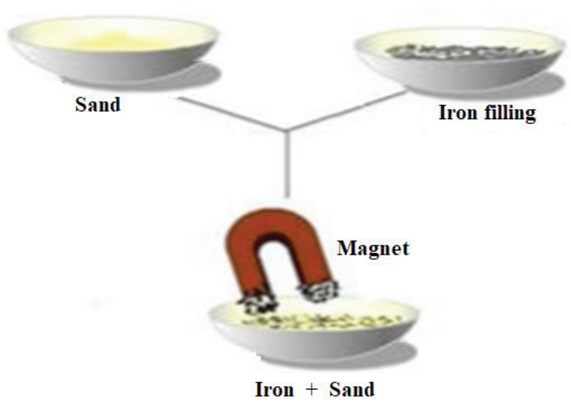
Chemicals

- Iron powder
- Petri dish/ plastic plate/bowl Plastic bag/wrapper
- Sand

Procedures

1. Mix the sand with the iron powder in the petri dish.
2. Wrap the plastic bag around the bar magnet.
3. Suspend the bar magnet over the plate.
4. Carefully remove the plastic bag around the magnetic bar and scrape off the iron powder.
5. Repeat the procedure from step 2 until all the iron powder are removed.

Experiment set-up



https://media.sciencephoto.com/image/c0435120/800wm/C0435120-Magnet_Pulls_Iron_Filings_From_Sand.jpg

Interpretation of results

Guiding/discussion questions

1. What type of mixture formed after mixing iron powder and sand?
2. What happens after putting the magnet above the mixture?
3. Why was the sand not attracted by the magnet?

Iron powder and sand form heterogeneous mixture. It is expected that iron powder will be immediately attracted to magnet bar leaving the sand behind.

Unit 7: Water and its composition

Experiment 7.1: Physical properties of water

Rationale: Water is a substance that has many uses in our daily life, it composes the biggest mass of all living organisms and serves as a medium for most of metabolic reaction, in chemistry it is mostly used as universal solvent for chemical compounds and also as a medium for chemical reactions.

This experiment is importantly used to show some physical properties of water like the difference between the boiling point and the melting point of pure water, it helps to know the boiling point and melting point of ice and other liquids by changing the temperature, density and electrical conductivity.

Objective: Learners will be able to determine the boiling point, melting point, density and electrical conductivity of pure water perfectly.

Required materials

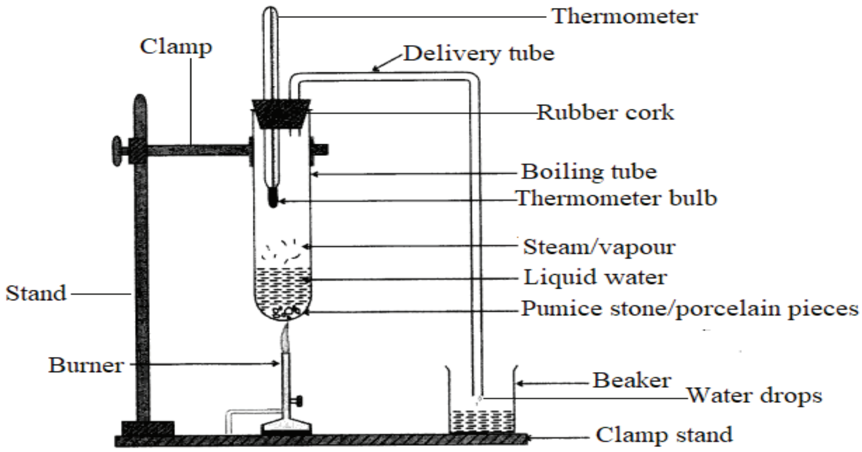
Apparatus

- Delivery tube
- Thermometer
- Bunsen burner
- Stand with a clamp and a boss
- Graduated beaker
- Boiling tube
- Rubber cork with two holes

Reagents

- Pumice stones
- Distilled water

Experiment set-up



Procedures

a) Investigating the boiling point of pure water

1. Take about 25ml of distilled water in a boiling tube and add 2-3 small pieces of pumice stone.
2. Close the mouth of the boiling tube with a rubber cork that has two bores and clamp it with the stand.
3. Introduce a thermometer (temperature range -10 to 110°C) in one bore of the cork of the boiling tube. Keep the bulb of the thermometer about 3-5 cm above the surface of the water.
4. Then introduce one end of a delivery tube in the second bore of the cork.
5. Place a 250ml beaker below the second end of the delivery tube to collect the condensed water.
6. Heat the boiling tube gently, preferably by rotating the flame.
7. Note the temperature (t_1) when the water starts boiling. Note the temperature by keeping your eyes in line with the level of mercury
8. Continue to heat the water till the temperature becomes constant, and the water remains boiling. Note the constant temperature (t_2)

Note: t_2 is the boiling point of water.

b) Investigating the melting point of pure water

Required materials

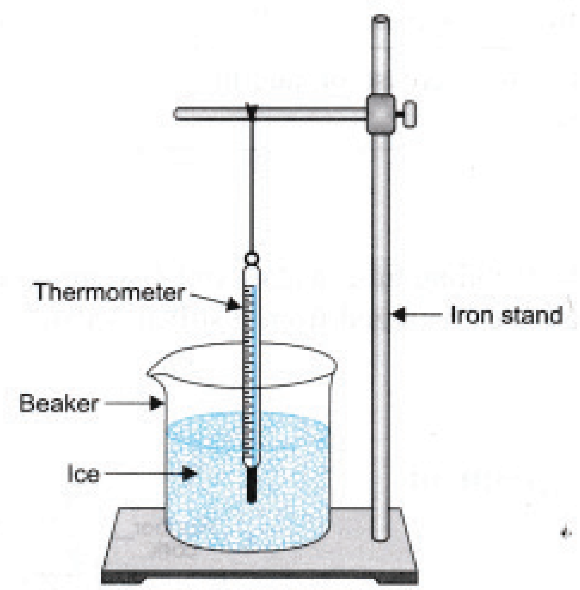
Apparatus

- Beaker
- Clamp
- Stand
- Thermometer (10°C-110°C)
- Gas burner or spirit lamp
- Tripod stand wire gauze
- Stop clock

Reagents

- Ice cream

Experiment set up



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Procedure

1. Take a beaker containing enough amount of ice,
2. Immerse the bulb of the thermometer in ice as shown here above,
3. Note the temperature (t1) of ice before starting heat it,
4. Heat the ice constantly and record the temperature after 1 minutes and until the time the all ice has melted,
5. Let record it (t2), note temperature by keeping your eyes in line with the level of mercury,
6. Do not remove the thermometer from the beaker until you record the readings.

Note: t2 is the melting point of water.

b) Determination of the density of water

Required materials

Apparatus

- Measuring cylinder,
- Dropper,
- Beaker,
- Weighing balance.

Reagents

- Distilled water.

Procedures

1. Weigh an empty measuring cylinder. Note down its mass (m1),
2. Pour 10 ml of distilled water in the measuring cylinder and reweigh Note down the mass (m2),
3. Find the mass of the distilled water by subtracting the first and second weighing: m2-m1,
4. Calculate the density by dividing the mass (m2-m1) by the volume (v),
5. Note: Density = (m2-m1): v.

Conclusion

Melting point of water	Boiling point of water	Density of water
0^0c	100^0c	$1g/cm^3$

Experiment 7.2: Test of the presence of water

Rationale: Water is a substance that can exist in all forms of physical states, serving universal solvent for most of other chemical substances, also as a common medium for chemical reactions.

This experiment helps to understand how to preserve some substances by keeping/storing them in the presence or absence of water. The experiment will also help us to understand the usefulness and harmfulness of the presence and absence of water to some substances with different properties

Objective: Learners will be able to test for the presence of water in a given substance.

Required materials

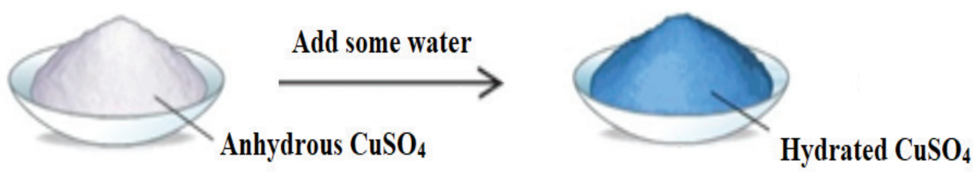
Apparatus

- Watch glass

Chemicals

- Anhydrous copper (II) sulphate
- Cold water

Experiment set up



Procedure

1. Put some anhydrous copper (II) sulphate on two separate watch glasses and label them as A and B.
2. Add some water drops to watch glass A and some pure ethanol drops to watch glass B.

Describe your observations.

Answer: The white powder on watch glass A turns blue whereas that on watch glass does not change.

Interpretation of results and conclusion

Guiding question:

1. Explain how water is tested using anhydrous copper (II) sulphate.

White anhydrous copper (II) sulphate becomes blue when mixed with some drops of water, but other liquids do not bring about the same change. Therefore, to test the presence of water anhydrous copper sulphate is used.

N.B: Water can also be detected using blue anhydrous cobalt (II) chloride (or a cobalt (II) chloride paper). This turns **pink** in the presence of water.

Unit 8: Air composition and pollution

Experiment 8.1: Determination of the percentage of oxygen in the atmosphere

Rationale: The primary source for the oxygen in the atmosphere is photosynthesis, where plants produce oxygen. The absence of the plants would increase the amount of carbon in the atmosphere and lower the amount of oxygen. Thus, deforestation alters the oxygen cycle. Mass deforestation has a neagative effect and it has therefore to be avoided.

Objective: Learners will be able to determine the percentage of oxygen in the atmosphere.

Required materials

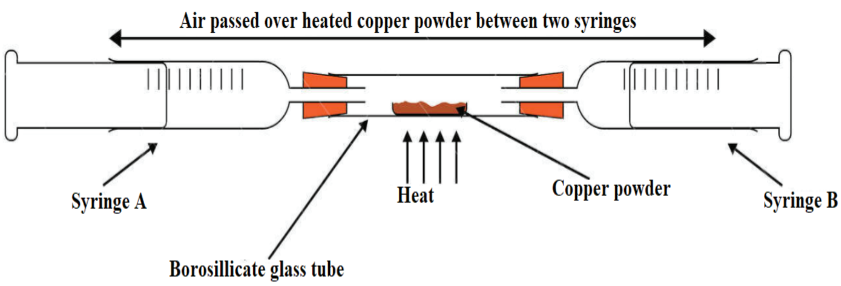
Apparatus

- 2 syringes
- Bunsen burner
- Hard glass tube
- Match box
- Retort stand and accessories
- Cork / stopper with one hole

Chemical

- Copper

Experiment set up



Procedure

1. Put a small amount of copper powder in a hard glass tube,
2. Fix the syringes on both ends of the glass tube,
3. Label the syringes as A and B,
4. Fill syringe A with 100cm3 of air by pulling out the plunger and remove all the air from syringe B by pushing inside the plunger,
5. Heat strongly the copper powder inside the tube and pass air over it by slowly pushing the plunger of the syringe A until there is no further change and allow the apparatus to cool down. Make sure all the air in syringe A is removed,
6. Record the volume of air in syringe B.

a) What is the colour of copper before heating?

Answer: *the colour of copper before heating is brown.*

b) What do you observe after heating?

Answer: *after heating the residue is black.*

Interpretation of results and Conclusion

Guiding question:

1. From the initial volume of air in syringe B and the final volume of air in syringe B, calculate the volume of air used in the experiment and its percentage by volume.

Brown copper burns in oxygen to form a grey black residue of copper (II) oxide.

Calculation of the volume of air used in the experiment

- Initial volume of air in syringe B = **100 cm^3**
- Final volume of air in syringe B = **79 cm^3**
- Volume of air used= 100 – 79 = **21 cm^3**
- Percentage of air used = (21:100) ×100 % = **21%**
- This means that only **21%** of air was used when copper was heated.

Oxygen is the active part of air. It is the only part of air which reacts with hot copper. It has **21 percentage** of the air by volume.

Unit 10: Chemical equations

Experiment 10.1: Verification of the law of conservation of matter

Rationale: How does the law of conservation of matter apply to our lives?

The law of conservation of matter and energy states that matter is neither created nor destroyed but conserved. Humans do not have the ability to create or destroy matter (atoms) or energy. They can only rearrange the matter and energy. For example, an oxygen atom will cycle through a living system

Objective

Learners will be able to verify and explain the law of conservation of matter in chemical reactions.

Required materials

Apparatus

- Measuring cylinders
- Conical flask
- Test tube
- Electronic balance

Chemical

- Barium chloride solution (0.1M)
- Sodium sulphate solution (0.1M)

Procedure

1. Measure 5 ml of sodium sulphate solution using a measuring cylinder and pour it into a conical flask.
2. Weigh the conical flask and its content and record their total mass, m_1.
3. Measure 5 ml of barium chloride solution using another measuring cylinder and pour it into the test tube.
4. Weigh the mass of the test tube and its content and record their mass, m_2.
5. Put the test tube in the conical flask and pour the solution from the test tube into the conical flask and mix well both solutions.

What do you observe?

Answer: *a white precipitate appears in the solution.*

6. Wait for ten minutes for the precipitate to settle at the bottom of the conical flask.
7. Weigh again the conical flask and its content and record their mass, m_3.

Data recording

Mass of the flask and sodium sulphate, m_1	Mass of the test tube and barium chloride, m_2	Mass of conical flask, test tube and their content, m_3

We note that $m_3 = m_1 + m_2$

Interpretation of results and Conclusion

Guiding question:

1. Explain how to verify the law of conservation of matter.

During chemical reactions, the starting substances (reactants) are changed into new substances (products). The changes that take place during chemical reactions are normally represented using chemical equations like the following hypothetical one

A + B → C + D, where A and B are reactants and C and D are products.

The equations are balanced to comply with the law of conservation of matter. This law states that matter cannot be created nor destroyed; but can only be transformed.

Equation of the reaction:

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

The total mass of the reactants is equal to the total mass of the products as predicted by the law of conservation of matter.

Evaluation activity

Write and balance the chemical equation of the reaction of calcium chloride and silver nitrate.

Unit 11: Acids and bases and pH

Experiment 11.1: Extraction of acid-base indicator from red onion leaves and flowers petals of Hibiscus

Rationale: This experiment permits to get natural non-toxic acid-base indicators ready to be used and enable to identify acids and bases.

Objective

Learners will be able to extract the indicators from leaves of plants or flower's petals.

Required materials

Apparatus

- Filter funnel
- Conical flask
- Test tubes
- Beaker
- Mortar and pestle
- Heat source
- Knife

Chemical

- Red onion/hibiscus flower
- Potassium hydroxide
- Dilute sulphuric acid
- Water

Procedure

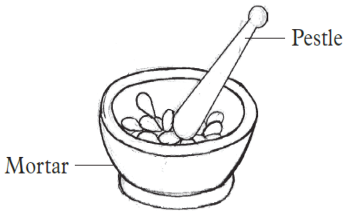
1. Cut off two leaves of red onion or hibiscus flowers into tiny pieces.
2. Crush the tiny pieces of red onion leaves or hibiscus flower in mortar using the pestle. Add boiling water as you continue crushing.

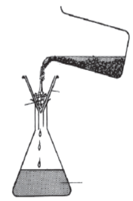
   Why do yu add boiling water?

   Answer: *This helps to extract as much dye from the leaves as possible.*
3. Allow the small pieces of red onion leaves to stand in the water until it becomes purple coloured. Transfer the contents into a beaker.

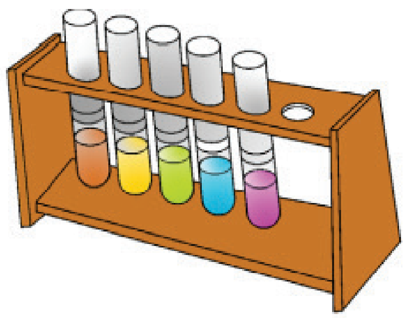
4. Filter the mixture into a clean conical flask. What is the colour of the filtrate?

5. Label clearly your extracts.





6. Add three drops of the red onion extract into five test tubes containing water, hydrochloric acid, sodium hydroxide, potassium hydroxide and dilute sulphuric acid respectively. Record your observations for each test tube.



Guiding questions:

1. What is the colour of red onion extract?
2. What is the colour of the juice extracted from hibiscus leaves?
3. What is the colour taken by solutions of HCl, H_2SO_4, NaOH and KOH when adding red onion extract and hibiscus leaves extract into each of them?
4. What is the role of boiled water in this experiment?

Conclusion

The red colour of onion extract turns pale red in acid and green in basic solution

The color of hibiscus extract in acidic solution is red and blue in basic solution.

Experiment 11.2: Testing different solutions using the common indictors

Rationale: The pH of water for drinking or for use in the home is very important. Water that is too alkaline or too acidic can damage pipes and appliances, and it is generally unhealthful to drink. The regulation of pH is very important in water treatment. Water naturally varies between about 6.5 and 8.5 on the pH scale, and this is normal.

Objective: Learners will be able to test different solutions using the common indicators

Required materials

Apparatus

- Test tubes
- Test tube racks
- Droppers
- Test tube holder

Chemicals

- Dilute Hydrochloric acid
- Dilute Sulphuric acid
- Sodium hydroxide solution
- Potassium hydroxide solution
- Water
- Ammonia solution
- Litmus papers
- Methyl orange
- Phenolphthalein
- Universal indicator

Procedure

1. Pour $5cm^3$ of dilute hydrochloric acid in four separate test tubes and label them as A, B, C and D.
2. Put a red litmus paper in test tube A, a blue litmus paper in test tube B, 2 drops of phenolphthalein indicator in test tube C and 2 drops of methyl orange in test tube D.
3. Record your observations in the table below.

Solution	**Indicator**	**Observations**
Dilute hydrochloric acid.	Red litmus	No change
	Blue litmus	Turns red
	Phenolphthalein	No change
	Methyl orange	Turns red

4. Repeat the experiment using dilute sulphuric acid, sodium hydroxide solution, potassium hydroxide solution, water, and ammonia solution separately.
5. Record your results in a table.

Solution tested	**Red litmus paper**	**Blue litmus paper**	**Phenolphthalein**	**Methyl orange**
Dilute Sulphuric acid	Stays red	Red	Colourless	Red
Sodium hydroxide solution	Blue	Stays blue	Pink	Yellow
Potassium hydroxide solution	Blue	Stays blue	Pink	Yellow
Water	Stays red	Stays blue	Colourless	Orange
Ammonia solution	Blue	Stays blue	Pink	Yellow

Interpretation of results and conclusion

Solution tested	Red litmus paper	Blue litmus paper	Phenolphthalein	Methyl orange
Acidic solution	Stays red	Turns red	Colourless	Red
Alkaline solution	Turns blue	Stays blue	Pink	Yellow
Neutral solution	Stays red	Stays blue	Colourless	Orange

Acid base indicators are substances that show different colors in acidic and alkaline media. They are used to test the acidity or alkalinity of a given medium.

Unit 12: Inorganic salts and their properties

Experiment 12.1: Solubility of inorganic salts

Rationale: The inorganic salts are important for living being, as cells use them to transmit electrical impulses across their membranes and to other cells throughout the body. The ions of inorganic salts (electrolytes) regulate nerve impulses,heart functions and muscle contractions.

Objective: Learners will be able to show whether salts are insoluble or soluble.

Required materials

Apparatus

- Test tube
- Spatula
- Bunsen burner
- Test tube holder
- Test tube rack
- Measuring cylinder

Chemicals

- Magnesium sulphate
- Calcium sulphate
- Sodium carbonate
- Lead (II) nitrate
- Potassium nitrate
- Silver chloride
- Sodium chloride
- Lead (II) chloride
- Barium sulphate

Procedure

1. Take one spatula full of a given salt and put it in test tube and label it by its formula. Put the test tubes containing various salts in a test tube rack,
2. Add some ml of water to the salt in each test tube and shake well for a while and return it into test tube rack,

3. Record the soluble and insoluble salts,
4. Take the test tubes containing insoluble salts with a test tube holder and heat them separately,
5. Record the salts that are soluble after heating and the ones which remain insoluble.

Observations

Salts	Soluble		Insoluble	
	Cold water	Hot water	Cold water	Hot water
Magnesium sulphate				
Calcium sulphate				
Sodium carbonate				
Lead (II) nitrate				
Potassium nitrate				
Silver chloride				
Sodium chloride				
Lead (II) chloride				
Barium sulphate				

Interpretation of results and conclusion

From observations made above complete the following table.

Soluble in cold water	Soluble in hot water	Insoluble
Sodium chloride	Lead (II) chloride	Silver chloride
Potassium nitrate		Barium sulphate
Sodium carbonate		
Lead (II) nitrate		
Magnesium sulphate		
Calcium sulphate		

Some salts are readily soluble in water while others dissolve only if the solution is heated. This is because generally the solubility of solids increases as temperature increases too. Salts which do not dissolve even after heating are said to be **insoluble**.

Note: All nitrates are soluble. All sodium, potassium and ammonium salts are soluble. All carbonate salts are insoluble except sodium carbonate, potassium carbonate and ammonium carbonate. All sulphate salts are soluble except lead (II) sulphate, barium sulphate, silver sulphate and calcium sulphate which is sparingly soluble. All chloride salts are soluble except lead (II) chloride and silver chloride. However, lead (II) chloride is soluble in hot water.

Experiment 12.2: Electrical conductivity of salts in aqueous solution.

Rationale: Conductivity is useful as a general measure of water quality. Each water body tends to have a relatively constant range of conductivity that, once established, can be used as a baseline for comparison with regular conductivity measurements.

Objective: Learner will be able to demonstrate and explain the electrical conductivity of salts in aqueous solution.

Required materials

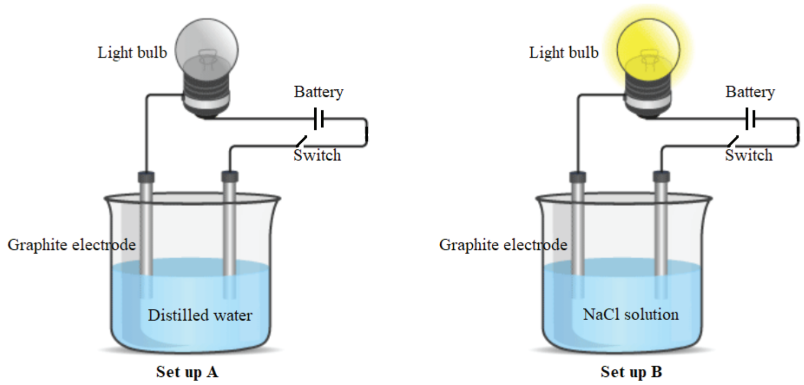
Apparatus

- Beakers or electrolysers
- Graphite electrodes
- Light bulb Sodium chloride
- Batteries Distilled water

Chemicals

- Magnesium sulphate
- Potassium nitrate

Experiment set up



Procedure

1. Set up the apparatus as shown on the figures A and B.
2. Separately test the electrical conductivity of distilled water and a solution of NaCl.
3. Switch on the current.

   What is your observation?

   Answer: *The bulb lights up when using a solution of NaCl, but it does not when using distilled water.*
4. Clean the electrodes with distilled water and repeat the experiment using magnesium sulphate and potassium nitrate solutions separately,
5. Record your observations in the table below.

Salt solutions/ Pure water	Observation on the bulb	Comment
Pure water	The bulb does not light up	Pure water does not conduct electricity
Magnesium sulphate	The bulb lights up	The solution conducts electricity
Potassium nitrate	The bulb lights up	The solution conducts electricity
Sodium chloride	The bulb lights up	The solution conducts electricity

Interpretation of results and Conclusion

Guiding questions

1. Explain why water does not conduct electricity while solutions of salts conduct electric current.

Pure water does not conduct electricity. All salts that dissolve in water form solutions that conduct electric current. This is because they dissolve to produce charged particles which move freely in the solution. These charged particles can carry the electric current. Generally, the electrical conductivity of a salt solution increases as the amount of dissolved salt increases.

Experiment 12.3: Thermal decomposition of salts

Rationale: Thermal decomposition of salts is a reaction that important in industries. For example, when calcium carbonate is heated it decomposes into calcium oxide and carbon dioxide. This process is employed in the manufacturing of quick lime, which is an important substance in many industries.

Objective

Learner will be able to carry out the thermal decomposition of various salts

Required materials

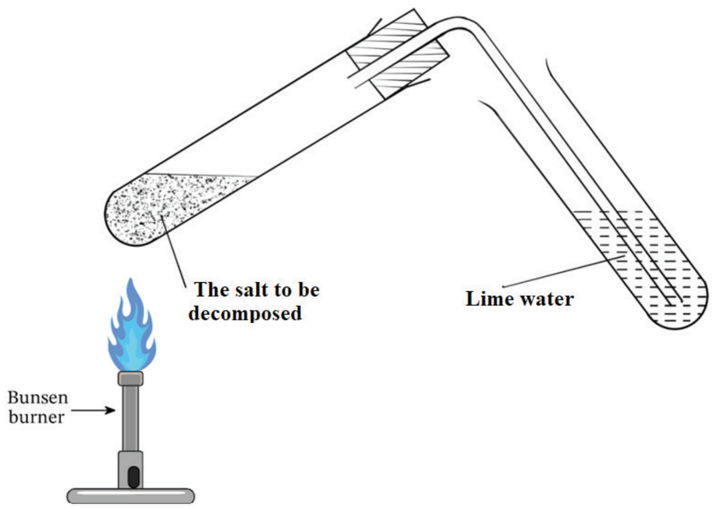
Apparatus

- Test tubes
- Spatula
- Bunsen burner
- Test tube holder
- Test tube rack
- Match box
- Delivery tube
- Cork with one hole

Chemicals

- Zinc carbonate
- Lead (II) carbonate
- Sodium carbonate
- Copper (II) carbonate
- Sodium hydrogen carbonate
- Calcium hydroxide solution
- Lead nitrate

Experiment set up



Procedure

1. Put into separate test tubes, about one spatula end-full of each of the provided chemicals,
2. Using a delivery tube, connect the test tube containing the salt to another test tube containing lime water,
3. Heat each sample gently and observe what happens,
4. Heat more strongly until no further change occurs,
5. Record your observations in the table bellow.

Observations

Salts	Colour before heating	Colour after heating	Effect of the gas on calcium hydroxide solution (lime water)
Zinc carbonate	White	Yellow when hot and white when cool	Lime water milky
Lead (II) carbonate	White	Red brown then Yellow	Lime water milky
Sodium carbonate	White	White	No observable change
Copper (II) carbonate	Blue green	Black	Lime water milky
Sodium hydrogen carbonate	White	White	Lime water milky
Lead nitrate	White	Red brown then Yellow	No observable change

Interpretation of results and Conclusion

Guiding questions

1. Which salts that do not decompose when heated and which ones are decomposed?
2. Identify the gas that makes lime water milky.

3. Write equations of reactions that occurs.
4. What is the brown red brown gas which becomes yellow when Lead nitrate decomposes?

Sodium carbonate does not decompose when heated because it is very stable.

Zinc carbonate decomposes on heating to form a yellow solid that turns white on cooling and a colorless gas is given off.

Lead (II) carbonate decomposes when heated to form a red brown solid that turns yellow on cooling while a black solid is formed when copper (II) carbonate is heated.

Sodium hydrogen carbonate decomposes on heating to give sodium carbonate, carbon dioxide and water.

The colorless gas evolved when a carbonate or hydrogen carbonate decomposes is Carbon dioxide. It turns lime water milky because of the formation of calcium carbonate $CaCO_3$ which is insoluble. However, if more carbon dioxide is bubbled through the milky solution, the white color disappears due to the formation of calcium hydrogen carbonate which is water soluble.

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

$$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2(s)$$

When lead nitrate is heated, brown fumes are released. The brown gas is nitrogen dioxide.

Equations of the reactions

1. $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
2. $ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$
3. $PbCO_3 \rightarrow PbO(s) + CO_2(g)$
4. $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$
5. $2Pb(NO_3)_2(s) \rightarrow 2PbO(g) + 4NO_2(g) + O_2(g)$

Unit 13: Preparation of oxygen and its properties

Experiment 13.1: Preparation of oxygen from hydrogen peroxide and its test.

Rationale: Oxygen is an essential medicine. Healthcare professionals use oxygen to treat respiratory illnesses like COVID-19 and pneumonia. Oxygen is also essential for surgery and trauma. Vulnerable groups like the elderly, pregnant women and newborns need oxygen in regular basis.

Objective: Learners will be able to prepare oxygen from some chemicals and test it.

Required materials

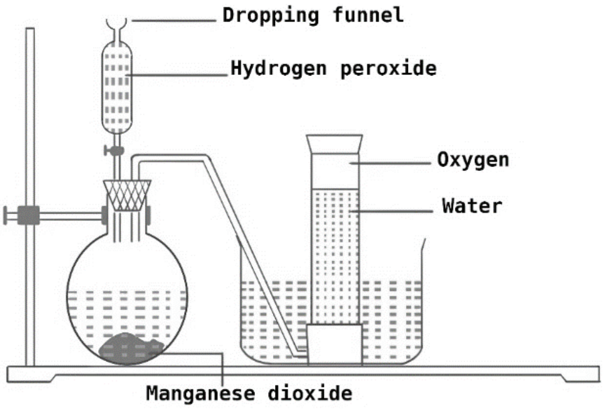
Apparatus

- Beehive shelf
- Electronic balance
- Stopper
- Delivery tube
- Flat-bottomed flask
- Dropping funnel
- Gas jar.
- Match box
- Trough

Chemicals

- Water
- Glowing splints
- Blue and red litmus papers
- Hydrogen peroxide
- Manganese dioxide

Experiment set up



Procedure

1. Weigh 3g of manganese dioxide using the weighing balance and put it in a flat-bottomed flask.
2. Set up the apparatus as shown.
3. Add hydrogen peroxide from the dropping funnel to the manganese dioxide in the flat-bottomed flask.

   What do you observe?

   Answer: A colorless gas which is collected over water is evolved.

   Is the gas soluble or insoluble in water? Explain your answer?

   Answer: The gas is insoluble in water because it passes through it and gets collected over it.
4. Test the gas collected using a glowing splint and the litmus papers.

   What do you observe?

   Answer: The gas relights a glowing splint. The gas is oxygen.
5. Test the effect of the gas on moist red or blue litmus paper.

   What do you observe?

   Answer: It has no effect on any wet litmus paper. It is a neutral gas.

Interpretation of results and conclusion

Guiding questions

1. What is the role of manganese dioxide in the experiment?
2. What is the natural gas which relights a glowing splint?
3. Which gas is collected by upward delivery over the surface of water and why?
4. Write equation of the reaction that occurs.

Hydrogen peroxide decomposes to produce oxygen. The decomposition of the peroxide is speeded up by manganese dioxide which acts as a catalyst. The oxygen gas produced has no effect on moist red or blue litmus paper. Oxygen is a neutral colorless and odorless gas. It relights a glowing splint and this serves as a chemical test for it.

Methods of collecting the gas depend on physical properties such as density and solubility of that gas. Oxygen is collected by upward delivery over the surface of water because it is less dense than air and insoluble in water.

Equation of the reaction:

$$2H_2O_2\,(aq) \rightarrow 2H_2O(l) + O_2(g)$$

Evaluation activity

What properties of the gas did you consider when collecting the gas?

Answer: We considered its density and solubility in water.

Experiment 13.2: Part of air that is used for burning

Rationale: Air contains about 21 percent oxygen, and most fires require at least 16 percent oxygen content to burn. Oxygen supports the chemical processes that occur during fire. When fuel burns, it reacts with oxygen from the surrounding air, releasing heat and generating combustion products. Effects of exposure to low oxygen concentrations can include giddiness, mental confusion, loss of judgment, loss of coordination, weakness, nausea, fainting, loss of consciousness and death.

Objective: Learners will be able to demonstrate that there is a part of air which is used in burning.

Required materials

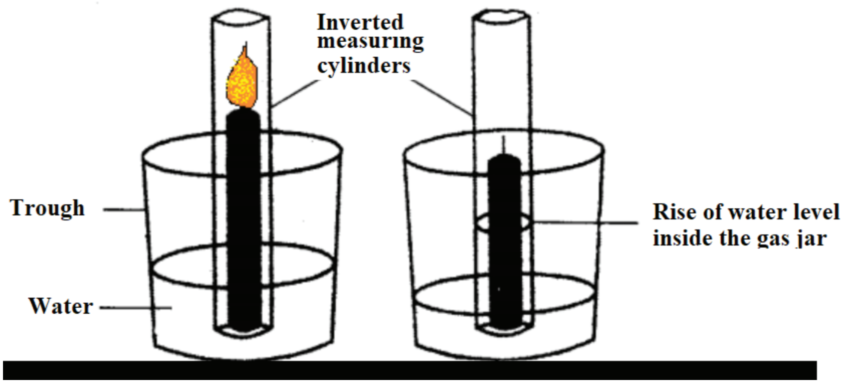
Apparatus

- Candle
- Measuring cylinders or gas jar
- Cork
- Matchbox
- Troughs

Chemicals

- Water

Experiment set up



Procedure

1. Put water in a trough as shown above,
2. Put a short candle on a cork and fix it in a trough containing water,
3. Measure the initial volume of the air inside the inverted measuring cylinder and record it as A,
4. Now remove the measuring cylinder and light the candle and cover it again gently with the same gas jar,

What do you observe?

Answer: The candle goes off and the level of the water in the measuring cylinder rises while the volume of air inside reduces.

Why does the volume of air inside the measuring cylinder reduces?

Answer: Part of the air is used up in burning.

5. Measure the final volume of air inside the measuring cylinder when the candle goes off and record it as B.
6. Using the results of this experiment, calculate the percentage of air used up in burning

Percentage of air used = (A-B/A) X100

Why is it necessary to let the apparatus to cool before the measuring of the final volume of the air inside the measuring cylinder?

Answer: Heating causes the expansion of gases

Interpretation of results and conclusion

Guiding questions

1. Why after burning for a while the candle goes off?
2. What is active part and inactive part of air?
3. Which gas is the active part of air and what is its percentage by volume of air?

After burning for a while, in a fixed volume of air, the candle goes off and the level of the water in the measuring cylinder rises. The candle goes off because the part of air used in burning is exhausted. The part of air used in burning is called the **"active part of air"** and therefore what remains in the gas jar is the **"inactive part of air"**. Since a part of the air is used up in burning, a partial vacuum is created in the gas jar. Greater atmospheric pressure acting on the surface of the water forces the level of water in the gas jar to go up.

The active part of air used in burning is **oxygen**. Oxygen forms about 21% by volume of air.

Experiment 13.3: Reactions of oxygen with metals

Rationale: When a metal reacts with oxygen, a metal oxide is formed. Metal oxides/ Mixed metal oxides have wide applications as catalyst, adsorbents, superconductors, semiconductors, ceramics, antifungal agents and also have spacious applications in medicines.

Objective

Learners will be able to carry out and explain the reactions of oxygen with metals.

Required materials

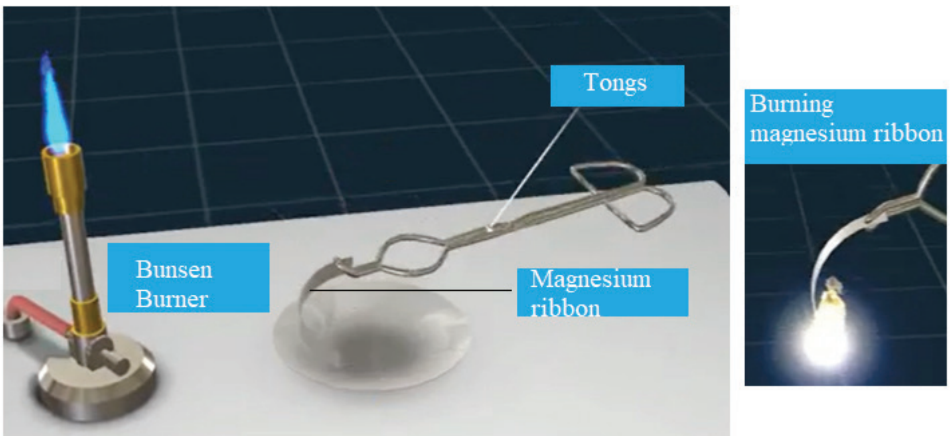
Apparatus

- Pair of tongs
- Bunsen burner
- Watch glass
- Match box

Chemicals

- Magnesium ribbon
- Calcium
- Iron
- Copper

Experiment set up



Caution

Avoid looking directly at the light source of the burning magnesium ribbon. It can cause temporally loss of sight.

Procedure

1. Take a small piece of magnesium ribbon, hold it with a pair of tongs and burn it.

   What do you observe?

   Answer: *The burning magnesium ribbon produces very bright white light and a white solid (magnesium oxide) is produced.*

2. Repeat the experiment for all other metals (calcium, iron and copper).

   What do you observe?

   Answer: *Calcium burns vigorously with a bright red flame in oxygen to form a white solid. When a red-hot iron wire is lowered in a gas jar of oxygen, it burns with yellow sparks. Copper burns in oxygen with a blue-green flame forming copper oxide that is a black powder.*

Interpretation of results and Conclusion

Guiding questions

1. Name the reactions that occur when oxygen reacts with metals (ex. Magnesium ribbon, Calcium, Iron and Copper) and write corresponding equations.
2. Name the product formed in each reaction carried out.

Magnesium, calcium, iron, and copper burn in oxygen in air to produce compounds called **oxides**. The reaction of oxygen and other elements or compounds is called **oxidation reaction**.

Equations of the reactions:

i. Magnesium burns brightly in oxygen to form a white powder, magnesium oxide.

   Magnesium + oxygen → magnesium oxide

   $$2Mg\ (s) + O_2(g) \rightarrow 2MgO(s)$$

ii. Calcium burns vigorously with a bright red flame in oxygen to form a white solid, calcium oxide.

Calcium + oxygen → calcium oxide

$$2Ca\ (s) + O_2(g) \rightarrow 2CaO(s)$$

iii. When a red-hot iron wire burns with yellow sparks to form oxide; iron (III) oxide.

Iron + oxygen → iron trioxide

$$4Fe\ (s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

iv. Copper burns in oxygen with a blue-green flame forming copper (II) oxide, which is black.

Copper + oxygen → copper (II) oxide

$$2Cu\ (s) + O_2(g) \rightarrow 2CuO(s)$$

Experiment 13.4: Reactions of oxygen with non-metals

Rational: Non- metals react with oxygen to form acidic oxides or neutral oxides. These oxides. These oxides are used, among others, to make gunpowder, fireworks, and matches to facilitate ignition. They are used in the manufacture of rubber for tires and other materials. They are also used as an insecticide or a fumigant.

Objective

Learners will be able to carry out the reactions of oxygen with non-metals

Required materials

Apparatus

- Deflagrating spoon
- Bunsen burner
- Watch glass
- Gas jar

Chemicals

- Sulphur
- Charcoal

Experiment set up

Procedure

1. Place 1 g of sulphur powder on a deflagrating spoon and heat it until it starts to burn. Lower the burning sulphur in a gas jar of oxygen.

   What happens?

   Answer: *Sulphur burns with a pale blue flame. It burns to form a colorless gas with a choking smell. Some white fumes are also formed.*

2. Place a small piece of charcoal on a deflagrating spoon and ignite it with fire until it is red-hot. Lower the burning charcoal in a gas jar of oxygen.

   What do you observe?

   Answer: *Charcoal reacts in the air at a glowing red heat to form a colorless gas which then burns in excess oxygen with a blue flame to produce a new gas.*

Interpretation of results and conclusion

Guiding question

1. Describe the reaction that occur when sulphur and charcoal burn in air and write corresponding equations.

Sulphur burns in oxygen with a pale blue flame. It burns to form a colorless gas with a choking smell. Some white fumes are also formed.

Sulphur burns in oxygen to form Sulphur dioxide.

$$S(g) + O_2(g) \rightarrow 2SO_2(g)$$

Charcoal reacts with oxygen at a glowing red heat to form a colorless carbon monoxide gas, which then burns with a blue flame with more oxygen from the air to produce carbon dioxide gas.

$$2C + O_2 (g) \rightarrow 2CO(g)$$

$$2CO + O_2(g) \rightarrow 2CO_2(g)$$

Experiment 13.5: Preparation of carbon dioxide

Rationale: In everyday life carbon dioxide is used as a refrigerant, in fire extinguishers, for inflating life rafts and life jackets, blasting coal, foaming rubber and plastics, promoting the growth of plants in greenhouses, immobilizing animals before slaughter, and in carbonated beverages

Objective: Learners will be able to prepare carbon dioxide and determine the method used for its collection.

Required materials

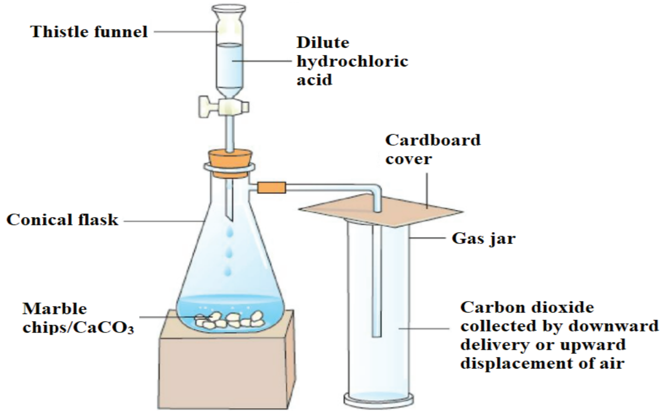
Apparatus

- Conical flask
- Thistle funnel
- Delivery tube
- 3 Gas jars with lids
- Cardboard cover

Chemicals

- Calcium carbonate
- Hydrochloric acid
- Limewater
- A lit splint
- Litmus paper

Experiment set up



Procedure

Preparation of the gas

1. Put some marble chips/ calcium carbonate into the conical flask,
2. Set up the apparatus as shown,
3. Put 50 cm^3 of hydrochloric acid in the thistle funnel,
4. Open the tap to let the acid flow into the conical flask,
5. Allow the gas jar to fill with gas until the reaction stops,
6. Repeat the steps 1-5 to get three jars filled with the gas,
7. Carry out the following tests.

Testing the gas

1. To the first gas jar put a lit splint.

   What do you observe?

   Answer: *the lit splint is extinguished.*

2. To the second gas jar add 1cm^3 of limewater, stopper, and shake.

   What happens?

   Answer: *the solution turns milky.*

3. To the third gas jar use wet litmus papers (red and blue).

   What do you notice?

   Answer: *moist blue litmus paper turns red, but the red does not change.*

Interpretation of results and conclusion

Guiding questions:

1. Explain the reactions that occur and write equations to explain these reactions.
2. Is CO_2 acidic or basic oxide and explain.

The reaction of calcium carbonate and hydrochloric acid produces calcium chloride, water, and carbon dioxide (a colorless and odorless gas).

Equation of the reaction

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

If a lit splint gets in contact with CO_2 it is extinguished. This is the reason why it used in fire extinguishers.

CO_2 also turns limewater (a solution of calcium hydroxide, $Ca(OH)_2$) milky and this is due to the formation of a white precipitate of calcium carbonate, $CaCO_3$.

Equation of the reaction

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

Being an acidic gas, CO_2 turns moist blue litmus red and has no effect on red litmus paper.

When collecting a gas, physical properties such as density and solubility in water should be considered.

Carbon dioxide is collected by downward delivery or upward displacement of air because it is denser than air.

Experiment 13.6: Preparation of hydrogen

Rationale: Hydrogen is used for fertilisers, grinding metals, and methanol to make ammonia for the manufacture of artificial materials such as plastics. Hydrogen is also used to create a strong explosion as a rocket fuel where liquid hydrogen is mixed with liquid oxygen.

Objective: Learners will be able to prepare and collect hydrogen.

Required materials

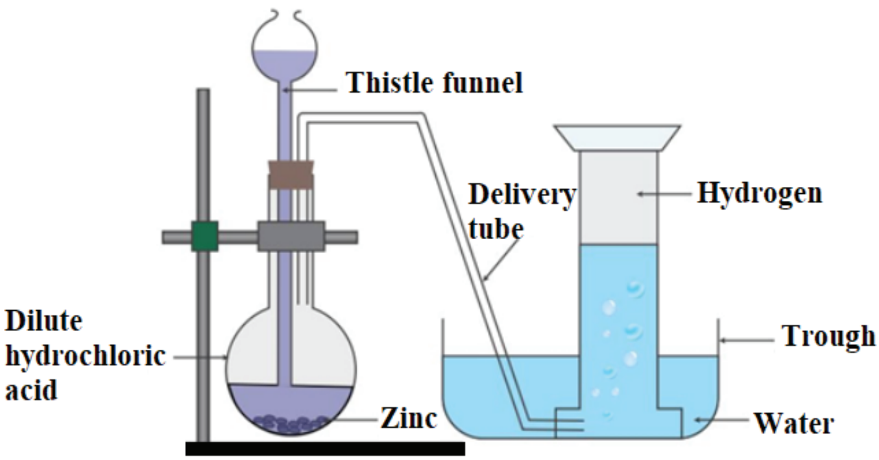
Apparatus

- Thistle funnel
- Delivery tube
- 3 gas jars with lids
- Water trough
- Electronic balance
- Round bottomed flask
- Stand and clamps
- Beehive shelf
- 3 Gas jars

Chemicals

- Zinc powder
- Dilute Hydrochloric acid
- Lit wooded splint
- Litmus paper

Experiment set up



Procedure

1. Measure 2.5g of zinc powder and put it into the round bottomed flask,
2. Set up the apparatus as shown and put 50 cm^3 of hydrochloric acid in the thistle funnel,

3. Open the funnel to let the acid flow into the flask and react with zinc,
4. Allow the gas jar to fill with gas until the reaction stops,
5. Repeat steps 1-4 to fill three gas jars,
6. Remove each gas jar from the water as required for the following tests.

Testing the gas

1. To the first gas jar put a lit splint.

   What do you observe?

   Answer: *Pop sound is produced.*

2. To the second gas jar add 1 cm^3 of limewater.

   What do you notice?

   Answer: *No observable change.*

3. To the third gas jar use both moist red and blue litmus papers.

   What happens?

   Answer: *No observable change.*

Interpretation of results and conclusion

Guiding question

1. Write equations of the reactions that occur and explain what is happening

Hydrogen gas is a colorless, odorless gas which is sparingly soluble in water.

When zinc reacts with hydrochloric acid, a colorless gas is produced. This gas that produces a pop sound in contact with a lit splint is **hydrogen gas.**

Equation of the reaction:

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$ZnCl_2$: zinc chloride

Some of the metals react with acids to produce hydrogen.

Example:

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Hydrogen gas has no effect on lime water and litmus papers. Thus, it is a neutral gas.

It is collected by upward delivery or downward displacement of water or air due to its low density and solubility in water.

Evaluation activity:

1. What properties of the gas did you consider when collecting it?

   Answer: *We consider its physical properties such as density and solubility in water.*

2. Complete and balance the following equations:

   a) $Fe(s) + HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

   b) $Ca(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$

Experiment 13.7: Preparation of sulphur dioxide

Rational: Sulphur dioxide in the preparation of sulphuric acid, sulphur trioxide, and sulphites, sulphur dioxide also is used as a disinfectant, a refrigerant, a reducing agent, a bleach, and a food preservative, especially in dried fruits.

Objective: Learners will be able to prepare sulphur dioxide and illustrate its method of collection.

Required materials

Apparatus

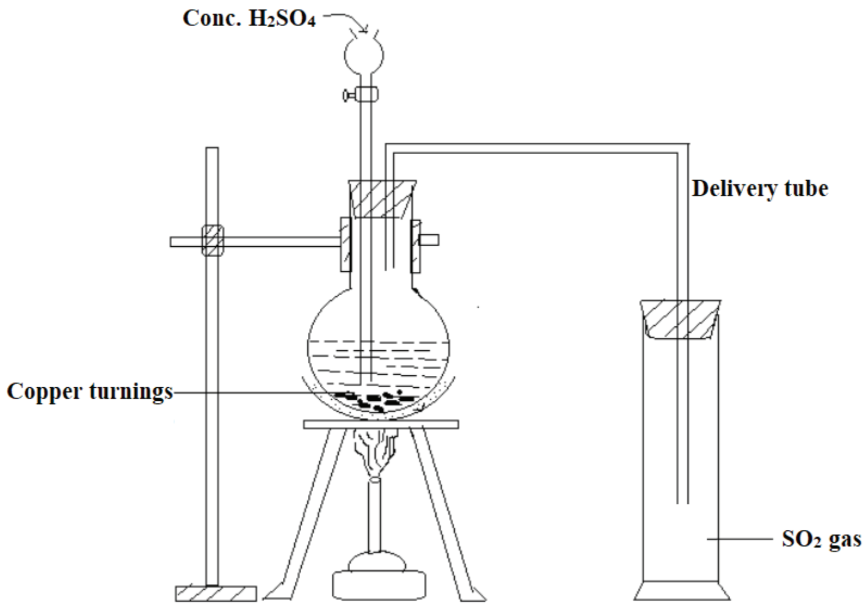
- Corks
- Thistle funnel
- Delivery tube
- 3 gas jars with lids

Chemicals

- Corks
- Thistle funnel
- Delivery tube 3 gas jars with lids
- Lit splint

- Bunsen burner
- Tripod stand
- Retort stand
- Round bottomed flask
- Wire gauze
- Electronic balance

Experiment set up



Procedure

Preparation of the gas

1. Measure 2g of copper using an electronic balance and put it into a round bottomed flask.
2. Set up the pieces of apparatus as shown on the figure.
3. Measure 50 cm^3 of concentrated sulphuric acid and put it in the thistle funnel.
4. Open the thistle funnel tap to allow the acid to flow into the flask it completely covers all copper turnings.

5. Gently heat the content pf the flask and let the gas jar get filled with the gas until there is no further change.
6. Repeat steps 1-5 to have three gas jars of gas.

Testing the gas

1. To the first gas jar insert a lit splint.

   What happens?

   Answer: *The gas in the jar extinguishes the splint.*

2. To the second gas jar add 1 cm^3 of acidified potassium dichromate solution. What do you notice?

   Answer: *The gas changes the color of acidified potassium dichromate solution from orange to green.*

3. Test the gas in the third jar using both wet red and blue litmus papers.

   What do you observe?

   Answer: *The gas turns moist blue litmus paper red, but it has no effect on red litmus paper. So, sulphur dioxide SO2 is an acidic gas.*

Interpretation of results and conclusion

Guiding questions:

1. Write a balanced equation to explain the preparation of sulphur dioxide
2. Is Sulphur dioxide acidic or basic?

Sulphur dioxide is usually made in the laboratory by reacting concentrated sulphuric acid with copper turnings.

Equation of the reaction:

$$Cu(s) + 2H_2SO_4(conc.) \rightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$$

Sulfur dioxide is a colorless chocking smelling gas which turns moist blue litmus paper and moist universal indicator paper red showing that it is acidic. It does not support burning and it changes the color of acidified potassium dichromate solution from orange to green. Therefore acidified potassium dichromate solution is used in identification test of sulfur dioxide

Being denser than air, sulfur dioxide gas is collected by downward delivery/ upward displacement of air or by using a gas syringe. It cannot be collected over water because it reacts with it to produce a sulphurous acid solution.

Note: The method of collecting gases using a gas syringe may be used to collect a gas of any density.

CHEMISTRY EXPERIMENTS FOR SENIOR TWO

UNIT 1 Chemical Bonding

Experiment 1.1: Physical properties of ionic and covalent compounds.

Rationale: In human life strong and weak bonds play key roles in the chemistry of our cells and bodies. For instance, strong covalent bonds hold together the chemical building blocks that make up a strand of DNA. However, weaker hydrogen bonds hold together the two strands of the DNA double helix.

In every day life, ionic compounds play important role, $NaCl$, sodium chloride ordinary table salt, NaF, sodium fluoride ingredient in toothpaste, $NaHCO_3$, sodium bicarbonate baking soda; used in cooking (and as antacid) …

Objective: Learners will be able to compare physical properties of ionic and covalent compounds (solubility, melting-boiling point and electric conductivity).

a) Solubility

Required materials

Apparatus

- Test-tube
- Test-tube rack
- Spatula

Chemicals

- Sodium chloride ($NaCl$)
- Water
- Ethanol
- Cooking oil

Procedure

1. Put a spatula endful of sodium chloride into a test tube;
2. Add 5 cm^3 of water and shake thoroughly;

   What happens?

   Answer: *Sodium chloride dissolves in water.*
3. Repeat the experiment with ethanol and water;

What happens?

Answer: *Ethanol dissolves in water.*

4. Repeat the experiment with cooking oil;

   What happens?

   Answer: *Cooking oil does not dissolve in water.*

Interpretation of results and conclusion

Ionic compounds like sodium chloride are soluble in water but in oil which is a covalent compound is not. Most ionic compounds are soluble in water whereas most covalent compounds are not. However, some covalent compounds such as hydrogen chloride are soluble in water.

b) Melting and boiling point

Required materials

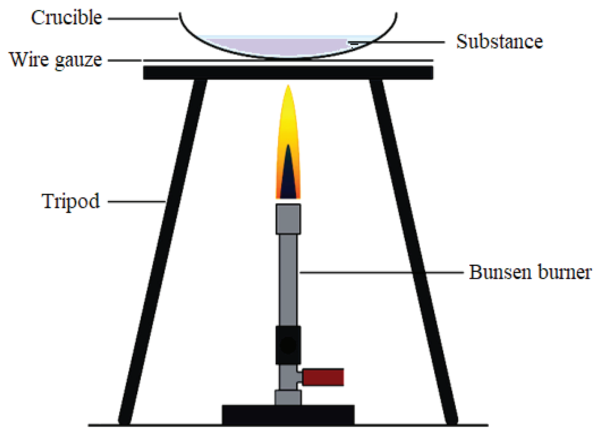
Apparatus

- Spatula
- Bunsen burner
- Retort stand and accessories
- Test-tube

Chemicals

- Candle wax
- Water
- Sodium chloride, (NaCl)

Experiment set up



Procedure

1. Place some common salt ($NaCl$) in a crucible;
2. Heat strongly for 5 minutes the crucible containing the salt as shown in the figure above.

   What do you observe?

   Answer: *After 5 minutes with strong heat, there is no observable change;*
3. Repeat the experiment by using candle wax;

   What do you observe?

   Answer: *After few seconds candle wax starts melting.*

Interpretation of results and conclusion

Guiding questions

1. Why sodium chloride needs strong heating while candle wax needs very short time to melt?

It requires strong heating and much time for sodium chloride to melt, while for candle wax in few seconds it starts to melt. Therefore, ionic compounds have high melting and boiling points while covalent compounds have low melting and boiling point. However, some covalent substances like graphite, diamond, silicon dioxide and aluminium oxide have high boiling and melting point.

c) Electric conductivity

Required materials

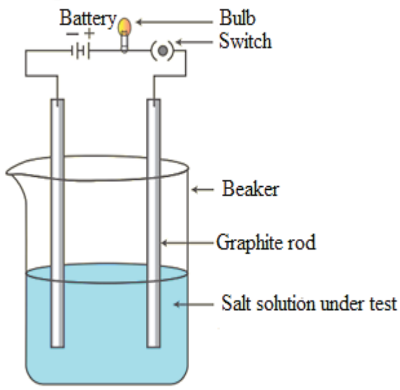
Apparatus

- Bulb
- Battery
- Switch
- Beakers
- Graphite rods

Chemicals

- Water
- Ethanol
- Sodium chloride ($NaCl$)

Experiment set up



Procedure

1. Place sodium chloride crystals in a beaker, add water and stir the mixture in two minutes;
2. Arrange the apparatus as shown in the figure above;
3. Switch on the electric current.

   What do you observe?

   Answer: *The bulb lights up.*
4. Repeat the experiment by using ethanol instead of salt solution.

   What do you observe?

   Answer: *The bulb does not light up.*

Interpretation

Guiding questions

1. Discuss and explain the electrical conductivity of covalent and ionic solutions.

For the case of sodium chloride solution, the bulb lights up means that it conducts electricity as it is ionic compound while for ethanol, the bulb does not light up meaning that it does not conduct electricity as it is covalent compound which does not dissociate into charges in water. Most covalent compounds do not conduct electricity while most ionic compounds conduct.

Unit 2: Trends in properties of elements in the periodic table

Experiment 2.1: Physical properties of metals and non-metals

Rationale: Metals are known to be good conductors of heat and electricity because they have free electrons. Some materials we use at home like saucepan, iron box, electric cables contain metals, metals play a big role in domestic economy. Electric cables are made of metals and therefore play a role in industrial economy.

Objective: Learners will be able to compare and contrast the physical properties of metals and non-metals

a) Thermal conductivity of metals and non-metals

Required materials

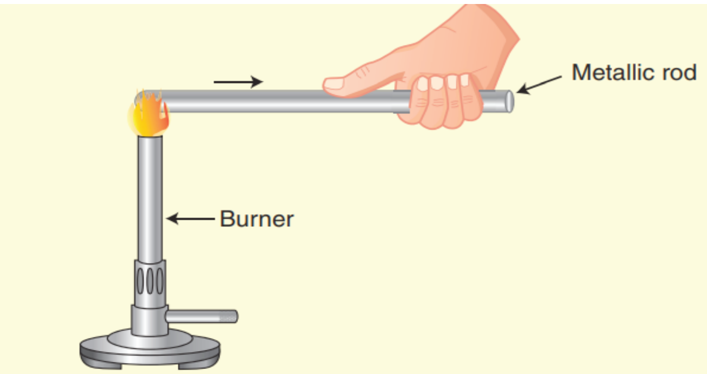
Apparatus

- Bunsen Burner
- Match box

Chemicals

- Metallic rod (iron rod)
- Sulphur rod

Experiment set up.



Conduction of heat in metals

Caution: Be careful while heating the objects, most of them are good conductors of heat

Procedure

1. Get a Bunsen burner and light it;
2. Take an iron rod, hold one end and put it to the flame of the burner for 3 to 4 minutes as it is shown in the experiment set up above;
3. Repeat the experiment by using sulphur rod instead of iron rod.

   What do you feel in those two experiments?

   Answer: *In the first experiment of heating iron rod, the hotness is felt while for the second experiment of heating sulphur rod, the hotness is not felt.*

Interpretation of results and conclusion

Guiding question

1. Explain why sulphur does not conduct heat while iron rod does.

In experiment one iron rod conduct heat as it is metal, while for the second experiment sulphur rod does not conduct heat as it is non-metal. Metals are good conductors of heat because they contain free electrons. These free electrons can move easily through the metal and conduct heat. Non-metals are not good conductors of heat because they do not contain free electrons.

b) Electrical conductivity of metals and non-metals

Required materials

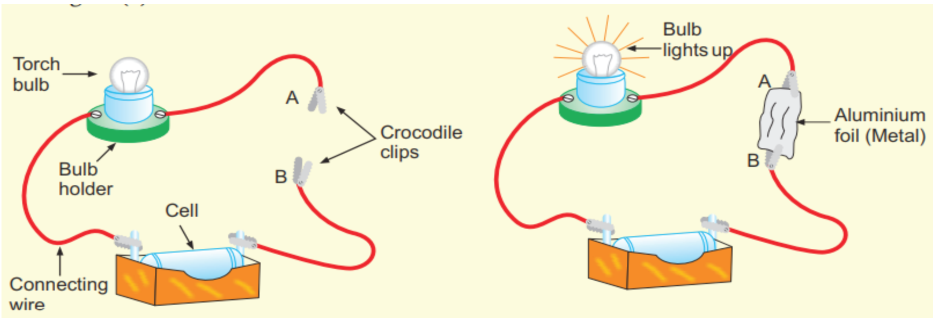
Apparatus

- Dry cell
- Torch bulb fitted in a holder
- Connecting wires
- Crocodile clips

Chemicals

- Aluminium foil/Iron rod
- Sulphur rod/Charcoal rod

Experiment set up



(*a*) (*b*)

Procedure

1. Take a dry cell, a torch bulb fitted in a holder and some connecting wires (ex. copper wires) with crocodile clips, and connect them as shown in diagram (a) to make an electric circuit;
2. Insert a piece of aluminium foil between the ends of crocodile clips A and B as shown in diagram (b).

   What do you observe?

   Answer: *When aluminium foil is inserted, the bulb lights up.*
3. Repeat the experiment by using sulphur or charcoal rod instead of aluminium foil;

   What do you observe?

   Answer: *When sulphur rod is used, the bulb did not light up.*

Interpretation of results and conclusion

Guiding question

1. Explain why sulphur does not conduct electricity while iron does.

The bulb lights up in case of aluminium foil as it is metal, but for the case of sulphur the bulb does not light up as it is non-metal.

Metals are good conductors of electricity because they contain free electrons. These free electrons can move easily through the metal and conduct electricity.

Most of non-metals are not good conductors of electricity because they do not contain free electrons.

Experiment 2.2: Reaction of metals with water, acids and oxygen.

Objective: Learners will be able to carry out, explain and compare the reaction of metals with water, acid and oxygen.

a) Reaction of metals with water

Required materials

Apparatus

- Beakers
- Stand and clamp
- Bunsen burner
- Match box
- Delivery tube
- Cork
- Boiling tube
- Trough
- Glass wool soaked in water

Chemicals

- Sodium (Na)
- Potassium (K)
- Calcium (Ca)
- Aluminium (Al)
- Iron (Fe)
- Magnesium (Mg)
- Zinc (Zn)
- Copper (Cu)
- Water

Caution: Do not touch sodium and potassium with bare hands, they cause severe burns.

Procedure

1. Put small pieces of sodium, potassium, calcium, aluminium, iron, magnesium, zinc and copper separately in beakers half filled with cold water;

   What do you observe?

   Answer: *When sodium and potassium are put in beakers, there is evolution of a gas, calcium floats after some times, while there is no observable change with other metals.*

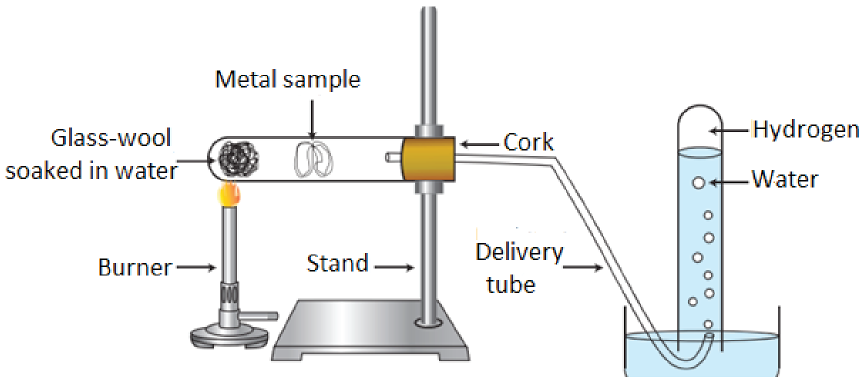
2. Put the metals that do not react with cold water in beakers filled with hot water.

   What do you observe?

   Answer: *Magnesium produces a gas with hot water.*

3. For those metals which did not react with hot water arrange the apparatus as shown in the figure below.

Experiment set up



What do you observe?

Answer: *Aluminium, Zinc and Iron react with steam and a gas is produced.*

Interpretation of results and conclusion

Guiding questions

1. Which products are formed when metals react with water?
2. Write equations of the reactions when K, Mg, Al, Zn, Fe and Ca and mention which metals react with cold water, hot water and vapors.

Some metals react with water to produce metal oxide and hydrogen gas. Sodium reacts vigorously with cold water to form sodium hydroxide and hydrogen gas.

$$\underset{\text{Potassium}}{2K(s)} + \underset{\text{Water (cold)}}{2H_2O(l)} \longrightarrow \underset{\text{Potassium hydroxide}}{2KOH(aq)} + \underset{\text{Hydrogen}}{H_2(g)} + \text{heat}$$

Potassium reacts violently with cold water to form potassium hydroxide and hydrogen gas.

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Water (cold)}}{2H_2O(l)} \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH(aq)} + \underset{\text{Hydrogen}}{H_2(g)} + \text{heat}$$

Calcium reacts with cold water to form calcium hydroxide and hydrogen gas.

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\text{Water (cold)}}{2H_2O(l)} \longrightarrow \underset{\text{Calciun hydroxide}}{Ca(OH)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Magnesium metal does not react with cold water. It reacts with both hot water and steam. It reacts with hot water to form magnesium hydroxide and water and very rapidly with steam to form magnesium oxide and hydrogen.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Water (hot)}}{2H_2O(l)} \longrightarrow \underset{\text{Magnesiun hydroxide}}{Mg(OH)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Aluminium reacts with steam to form aluminium oxide and hydrogen gas.

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\text{Steam}}{3H_2O(g)} \longrightarrow \underset{\text{Aluminium oxide}}{Al_2O_3(s)} + \underset{\text{Hydrogen}}{3H_2(g)}$$

Zinc reacts with steam to form zinc oxide and hydrogen gas.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Steam}}{H_2O(g)} \longrightarrow \underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Iron reacts with steam to form iron oxide and hydrogen gas.

$$\underset{\text{Iron}}{3Fe(s)} + \underset{\text{Steam}}{4H_2O(g)} \longrightarrow \underset{\text{Iron oxide}}{Fe_3O_4(s)} + \underset{\text{Hydrogen}}{4H_2(g)}$$

Copper does not react with water (or steam).

$$\underset{\text{Copper}}{Cu(s)} + \underset{\text{Water or Steam}}{H_2O(l/g)} \longrightarrow \text{No reaction}$$

b) Reaction of metals with acids.

Required materials

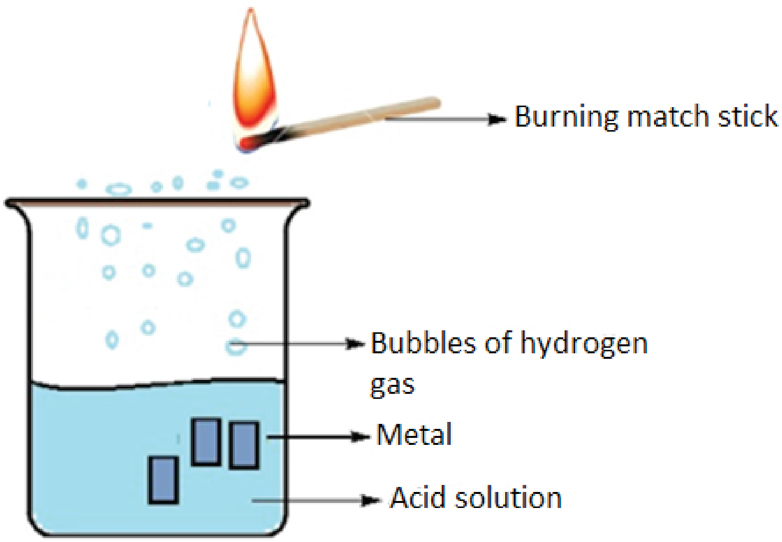
Apparatus

- Beaker

Chemicals

- Hydrochloric acid, (HCl)
- Magnesium, (Mg)

Experiment set up.



Procedure

1. Put a piece of magnesium into a beaker containing 50 cm^3 of 1M hydrochloric acid;
2. Take a light matchstick closer to the mouth of the tube containing the product of the reaction.

   What do you observe?

   Answer: *Hear a pop sound. When the magnesium stops reacting and no further bubbles are released, reaction is over.*

Interpretation

Guiding question

1. Describe the reaction that occur when metals react with acids in general and when magnesium reacts with hydrochloric acid.

Bubbles formed on the surface of the magnesium pieces means hydrogen gas is produced. Some metals react with acid to produce metal salt and hydrogen gas. Magnesium reacts with dilute hydrochloric acid to give magnesium chloride and hydrogen gas.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Dilute hydrochloric acid}}{2HCl(aq)} \longrightarrow \underset{\text{Magnesium chloride}}{MgCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

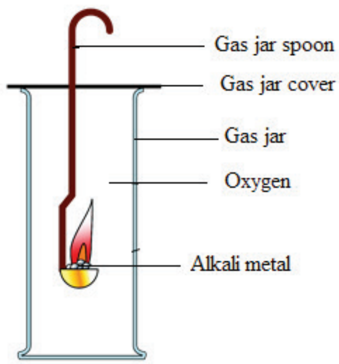
c) Reaction of metals with oxygen

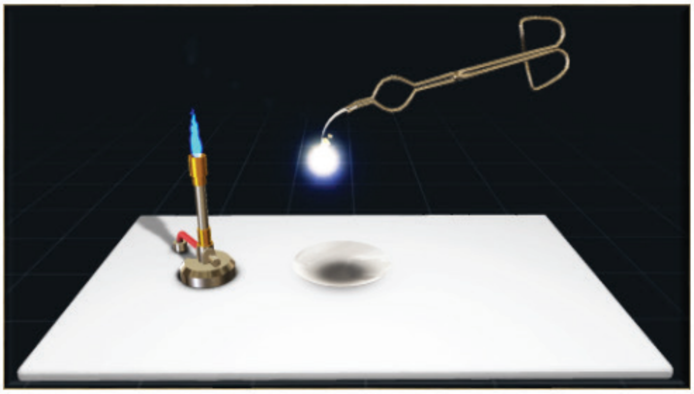
Required materials

Apparatus
- Gas jar
- Gas jar spoon/deflagrating spoon
- Gas jar cover

Chemicals
- Magnesium
- Oxygen

Experiment set up.





Procedure

1. Fill gas jars with oxygen;
2. Heat metal on gas jar over a flame;
3. Place the heated element in a cylinder of oxygen and observe the flame colour;

   What do you observe?

   Answer: *Bright white flame is observed when Magnesium is ignited as shown in the diagram above right.*

Interpretation

Magnesium burns brightly in oxygen to form a white powder; magnesium oxide.

$$\underset{\text{Magnesium}}{2Mg(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Magnesium oxide}}{2MgO(s)}$$

Metals react with oxygen to form metal oxides but it depends of reactivity series of metals, for instance, potassium and sodium reacts very vigorously with oxygen at room temperature and catch fire but those less reactive need heat to react with oxygen.

Unit 5: Categories of chemical reactions

Experiment 5.1: Reaction of hydrogen chloride with ammonia

Objective: Learners will be able to show the reaction between ammonia and hydrogen chloride.

Required materials

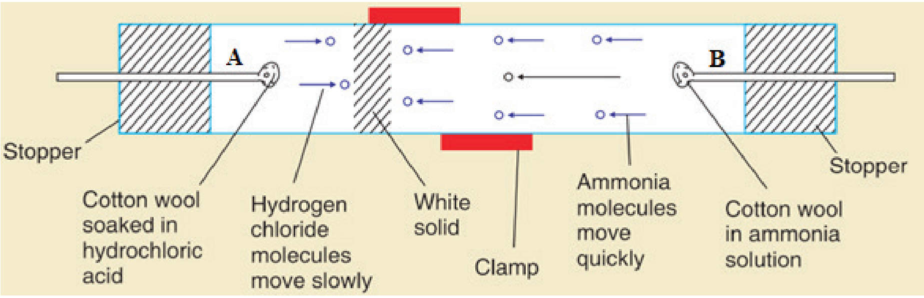
Apparatus

- Glass rods (two)
- Cotton buds
- Long glass tube
- Stoppers (two)
- Clamp

Chemicals

- Ammonia solution, NH_4OH (aq)
- Hydrochloric acid, HCl(aq)

Experiment set up



Caution:

Put on hand gloves before conducting experiment, clamp glass tube properly and mask nose and mouth.

Ammonia and hydrogen chloride gases are poisonous. Point the bottle away from both you and the students.

Procedure

1. Fix a cotton bud on two glass rods;
2. Mark the glass rods as A and B;
3. Dip the rod A into concentrated ammonia solution and rod B into concentrated hydrochloric acid;

   What do you observe?

   Answer: *When ammonia reacts with hydrochloric acid, a white solid appears inside the glass tube.*

Interpretation

In this experiment, ammonia reacts with hydrochloric acid and white fumes of ammonium chloride (NH_4Cl) are formed inside the glass tube. Here, two reactants NH_3 and HCl react to give one product, NH_4Cl. This type of reaction is called combination reaction.

Equation of the reaction: $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

Experiment 5.2: Decomposition of hydrated copper (II) sulphate

Objective: Learners will be able to show that hydrated copper (II) sulphate decomposes by heat.

Required materials

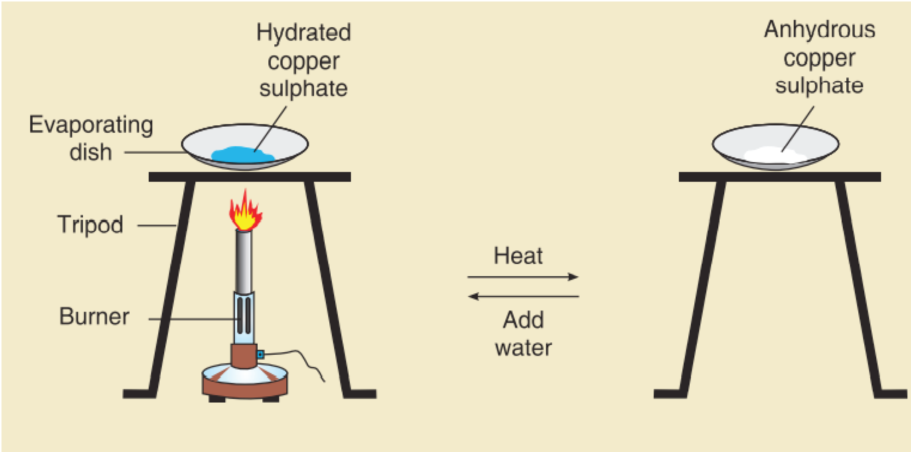
Apparatus

- Evaporating dish
- Tripod stand
- Bunsen Burner
- Balance
- Wire gauze.

Chemicals

- Copper sulphate crystals

Experiment set up



Caution: Put on safety glasses, do not look directly into the evaporating dish, keep away from the burner; do not touch the hot dish.

Procedure

1. Put 2 g of copper sulphate crystals in an evaporating dish;
2. Fix all apparatus as shown in the figure above;
3. Heat hydrated copper sulphate gently.

   What do you observe?

   Answer: *Blue copper sulphate crystals become white when heated.*
4. Add water to the white copper sulphate and note your observation.

   Observation: *The white powder becomes blue again*

Interpretation

Guiding questions

1. What is the effect of heat on the hydrated copper sulphate (blue)?
2. What do you observe when water is added to anhydrous copper sulphate?
3. Write a balanced equation of the effect of heat on hydrate copper sulphate.

The heat causes to split into anhydrous copper sulphate and water. Anhydrous copper sulphate is white in colour. If we add water to the anhydrous copper sulphate, the white powder becomes blue again.

Equation of the reaction:

$$\underset{\text{Hydrated copper su;phate (blue)}}{CuSO_4.5H_2O} \underset{}{\overset{\text{Heat}}{\rightleftharpoons}} \underset{\text{Anhydrous copper sulphate (white)}}{CuSO_4} + \underset{\text{Water}}{5H_2O}$$

The decomposition reactions require energy either in the form of heat, light or electricity for breaking down the reactants.

Other compounds like calcium carbonate can also undergo decomposition reaction.

$$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

Experiment 5.3: Precipitation reaction

Rationale: In medicine precipitation reaction is a test involved in the serology, for the qualitative and quantitative detection of antigens and antibodies. It is a reaction, where the soluble reactants (antigen and antibody) convert into an aggregated form.

Objective: Learners will be able to show and explain precipitation reactions.

Required materials

Apparatus

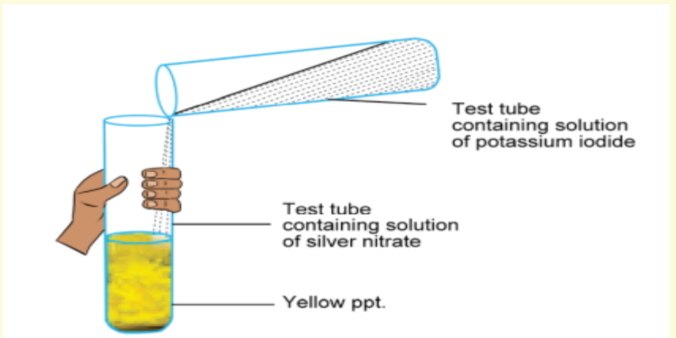
– Test tube

Chemicals

– Silver nitrate solution

– Potassium iodide solution.

Experiment set up



Formation of silver iodide

Procedure

Take silver nitrate solution in a test tube and add potassium iodide solution as it is shown in the diagram above.

What do you observe?

Answer: Yellow precipitate is formed.

Interpretation of results and conclusion

Guiding questions

1. What do you observe when silver nitrate solution with potassium iodide solution?
2. What is the yellow precipitate?
3. Name the reaction between silver nitrate with potassium iodide.
4. Write equation of the reaction between silver nitrate with potassium iodide.

When silver nitrate solution reacts with potassium iodide solution, a yellow precipitate is formed. This precipitation reaction is a double displacement reaction between silver nitrate and potassium iodide.

$$AgNO_3(aq) + KI(aq) \rightarrow AgI(s) + KNO_3(aq)$$

In this double displacement reaction, two compounds, silver nitrate and potassium iodide, react by an exchange of ions to form two new compounds, silver iodide and potassium nitrate.

Evaluation activity

1. Perform experiment to show the reaction between diluted sulphuric acid and barium nitrate.

   a) Is this reaction a precipitation reaction? Explain

   b) Write equation of reaction that occurs.

   c) Write the chemical formula of the precipitate.

Experiment 5.4: Differentiation of endothermic and exothermic reaction

Rationale: Exothermic reaction is a reaction that products energy in the form of heat. Many exothermic reactions help us in our daily life. The burning of fuel is an example of a combustion reaction and human needs this process for his energy requirements. The following equations describe the combustion of a hydrocarbon such as petrol. The chemical reaction that takes place when fuels burn has both positive and negative consequences. Although we benefit from heat, the carbon dioxide that is produced has a negative impact on the environment. Photosynthesis is an endothermic reaction. This means it cannot occur without energy (from the Sun). The light required is absorbed by a green pigment called chlorophyll in the leaves.

Objective: Learners will be able to differentiate exothermic and endothermic reactions.

Required materials

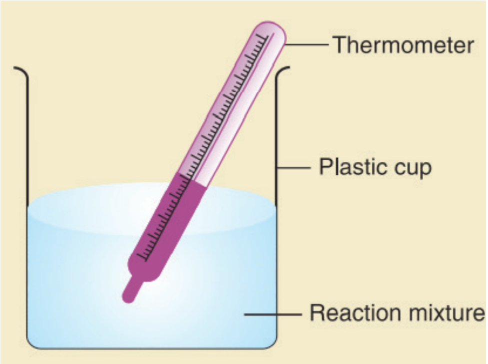
Apparatus

- Beaker
- Plastic cup
- Thermometer

Chemicals

- 10 cm^3 of sodium hydroxide solution
- 10 cm^3 of dilute hydrochloric acid.

Experiment set up



Caution: Put eye glass because some of the solutions are irritant and use the thermometer properly to measure changes in temperature.

Procedure

1. Put 10 cm^3 of sodium hydroxide solution in the beaker and record the temperature;
2. Add 10 cm^3 of dilute hydrochloric acid, stirring with the thermometer;
3. Record the temperature;
4. Repeat the procedure by using:

   a) Sodium hydrogencarbonate solution and citric acid;

   b) Copper (II) sulphate solution and magnesium powder;

   c) Dilute sulphuric acid and magnesium ribbon.

Data recording

Reaction	T° before reaction (t_1)	T° after reaction (t_2)	Exothermic or Endothermic reaction
Sodium hydroxide solution + dilute hydrochloric acid			
Sodium hydrogencarbonate solution + citric acid			

Copper (II) sulphate solution + Magnesium powder			
Dilute sulphuric acid + Magnesium ribbon			

Interpretation of results and conclusion

Guiding questions

1. What is exothermic and endothermic reaction?
2. Give same examples of endothermic and exothermic reactions.

The chemical reactions which proceed with the evolution of heat energy are called **exothermic reactions** i.e. the temperature after the reaction is higher than the temperature before the reaction. The chemical reactions which proceed with the absorption of heat energy are called **endothermic reactions** i.e. the temperature after the reaction is lower than the temperature before the reaction. Examples of endothermic reactions include the reaction of barium hydroxide with ammonium chloride, the thermal decomposition of calcium carbonate, and photosynthesis.

Experiment 5.5: Displacement of copper from its solution by zinc

Rationale: In displacement reactions, a more reactive element can displace a less reactive one out of its compound during a chemical reaction. Displacement reactions can be used to investigate the reactivity of metals and extract metals from metal oxides. In real life some applications of displacement reactions are thermite welding, steel making, extraction of metals, and relief from acid indigestion.

Objective: Learners will be able to show how copper is displaced from its solution.

Required materials

Apparatus

– Beaker

Chemicals

– Copper (II) sulphate solution

– Zinc metal

Experiment set-up

Before the reaction

After the reaction

Procedure

1. Put 10 mL of copper (II) sulphate solution in a beaker;

   What is the colour of the solution in the beaker?

   Answer: *Blue*

2. Dip a piece of zinc metal in the beaker.

   What do you observe?

   Answer: *After few minutes, a reddish-brown solid is formed on the surface of the zinc strip and the blue solution changes into colourless.*

Interpretation

Guiding questions

1. When zinc is left in copper (II) sulphate solution, it gradually decay. Explain this.
2. Write equation of the reaction that occurs.
3. Between zinc and copper which metal is more reactive? Why?

When a strip of zinc metal is placed into a blue solution of copper (II) sulphate a reaction immediately begins as the zinc strip begins to darken. If left in the solution for a longer period of time, the zinc will gradually decay due to oxidation to zinc ions. At the same time, the copper (II) ions from the solution are reduced to copper metal which causes the blue copper (II) sulphate solution to become colourless.

In the reaction zinc as a more reactive metal has displaced copper in copper (II) sulphate. This means that more reactive metal replaces less reactive metal from its salt solution.

Equation of the reaction:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Unit 6: Preparation of salts and identification of ions

Experiment 6.1: Effect of temperature on solubility of different salts

Objective: Learners will be able to show the effect of temperature on the solubility of different salts.

Required materials

Apparatus

- Beakers
- Electronic balance

Chemicals

- Water
- Potassium chloride
- Potassium nitrate
- Calcium carbonate
- Lead (II) chloride

Procedure

1. Put 100 g of water in each of four separate beakers at room temperature;
2. Add potassium chloride to the first, potassium nitrate to the second, calcium carbonate to the third and lead chloride to the fourth beaker;
3. Stir well to dissolve the required solute; and more solute until no more salt can be dissolved. You have got a "saturated solution".
4. Heat separately each solution up to 60°C and add more solute until you get again a saturated solution.

   What do you observe when the solution is heated?

   Answer: *Water can dissolve more solute.*

Interpretation of results and conclusion

Guiding questions

1. What is the effect of temperature on solubility?

The results of this experiment show that for different substances, the amount of solute needed to make a saturated solution in the same solvent is different at a particular temperature. The amount of solute (in grams) required to form a saturated solution in 100 grams of solvent (water) at a particular temperature is called the solubility of the substance at that temperature.

Generally, the solubility of the solute increases with the increase in temperature.

Experiment 6.2: Preparation of Copper (II) sulphate from copper (II) oxide and dilute sulphuric acid

Rationale: Copper sulfate can be prepared by reacting oxide of copper with dilute sulphuric acid. Copper sulfate has many important applications in different domains of real life. It is used as a fungicide, root killer, and herbicide in both agriculture and non-agricultural settings. Copper sulfate is used as a fungicide, algaecide, root killer, and herbicide in both agriculture and non-agricultural settings. It is also used as an antimicrobial.

Objective: Learners will be able to prepare copper (II) sulphate from copper (II) oxide and dilute sulphuric acid.

Required materials

Apparatus

- 250 cm^3 beaker
- Bunsen burner
- Evaporation dish
- Filter paper

Chemicals

- 50 cm^3 of dilute H_2SO_4
- Copper (II) oxide

Procedure

1. Take 50 cm^3 of dilute H_2SO_4 in a 250 cm^3 beaker and warm it gently;
2. Add black copper (II) oxide to the above acid in small portions until there is no more oxide that dissolves;
3. Filter the solution to remove unreacted copper (II) oxide;
4. Heat the bluish filtrate to concentrate in an evaporation dish;
5. Allow the solution to cool undisturbed.

   What do you observe?

   Answer: *Blue crystals of $CuSO_4$ (blue vitriol) are formed after cooling.*

Interpretation of results and conclusion

Guiding questions

1. What do you observe when copper (II) oxide reacts with dilute sulphuric acid?
2. Write equation of the reaction that occurs.

When copper (II) oxide reacts with dilute sulphuric acid, a blue crystal of copper sulphate is formed.

Equation of the reaction:

$$CuO\ (s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

Experiment 6.3: Preparation of zinc sulphate crystals from zinc metal and diluted sulphuric acid

Objective: Learners will be able to prepare zinc sulphate crystals from zinc metal and diluted sulphuric acid.

Required materials

Apparatus

- 250 cm^3 beaker
- Evaporating dish
- Filter paper

Chemicals

- 40 cm^3 of dilute H_2SO_4
- 10 g of zinc granules

Procedure

1. Put 40 cm^3 of dilute H_2SO_4 in a beaker and add 10 g of zinc granules;
2. Record your observations;

   What do you observe?

   Answer: *Evolution of gas and after a while evolution of gas ceases because the reaction is over.*
3. Once the reaction is over, filter the solution to remove unreacted zinc;
4. Heat the obtained solution in a dish;
5. Allow solution to cool and concentrate to crystallization point;
6. Separate and dry in the folds of a filter paper.

   What do you observe?

   Answer: *The needle shaped crystals of white vitriol are obtained after cooling.*

Interpretation of results and conclusion

Guiding questions

1. How is zinc sulphate prepared?
2. Write equation of the preparation of zinc sulphate showing physical states reactants and products.

Zinc sulphate is a water-soluble salt. It can be prepared in the laboratory by the action of dilute H_2SO_4 on zinc granules.

Equation of reaction:

$$Zn\ (s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2\ (g)$$

Experiment 6.4: Preparation of calcium chloride from calcium carbonate and hydrochloric acid

Rationale: Calcium chloride is prepared from calcium carbonate and hydrochloric acid. It is a prescription medicine used to treat the symptoms of hypocalcaemia (too little calcium in your blood), arrhythmias (problem with the rate or rhythm of your heartbeat), hypomagnesemia (a low level of magnesium in the blood). Calcium Chloride may be used alone or with other medications.

Objective: Learners will be able to prepare calcium chloride from calcium carbonate and hydrochloric acid.

Required materials

Apparatus

- Measuring cylinder (250 cm^3)
- Beakers
- Evaporating dish
- Glass spatula
- Filter funnel
- Filter paper
- Electronic balance

Chemicals

- Hydrochloric acid (50 cm^3)
- Calcium carbonate powder (6g)

Procedure

1. Put 50 cm^3 dilute hydrochloric acid in a beaker;
2. Weigh 6.0 g of calcium carbonate by using an electronic top pan balance and put it in a small plastic bottle;
3. Take a spatula of calcium carbonate from the plastic bottle into the acid in the beaker, stir the mixture.

   What do you observe?

   Answer: *Calcium carbonate dissolves.*
4. Continue to add more calcium carbonate to the acid.

What are your observations?

Answer: *When continuing adding more calcium carbonate, the dissolution stops and some the solid is left at the bottom of the beaker.*

5. Filter the excess calcium carbonate with a piece of filter paper and a funnel;
6. Collect the filtrate and place it into an evaporating dish.

What do you observe?

Answer: *After evaporating, crystals of calcium chloride are obtained.*

Interpretation of results and conclusion

Guiding questions

1. How is calcium chloride prepared?
2. Write equation of the reaction of the preparation of calcium chloride.

When calcium carbonate reacts with hydrochloric acid, calcium chloride is formed and carbon dioxide is released.

Equation of the reaction:

$$CaCO_3(s) + HCl\ (aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Experiment 6.5: Preparation of lead carbonate from lead nitrate and sodium carbonate

Objective: Learners will be able to prepare lead carbonate from lead nitrate and sodium carbonate.

Required materials

Apparatus

- Test tubes
- Glass rod

Chemicals

- Sodium carbonate
- Lead nitrate

Procedure

1. Prepare aqueous solutions of sodium carbonate and lead nitrate with water in two different beakers;
2. Take 20 mL of sodium carbonate solution in a test tube;
3. Add 20 mL lead nitrate solution to the test tube;

What do you observe?

Answer: *White precipitate (of lead carbonate) is produced.*

Interpretation of results and conclusion

Guiding questions

1. How is lead carbonate prepared from lead nitrate?
2. Write equation of the reaction of this preparation.

The lead nitrate and sodium carbonate have exchanged their ions. This is an example of double replacement reaction.

Equation of reaction:

$$Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2NaNO_3(aq).$$

Experiment 6.6: Identification of cations and anions

Rationale: The most abundant cations present in water are calcium (Ca^{2+}), magnesium (Mg^{2+}), sodium (Na^{+}), and potassium (K^{+}); the most abundant anions are bicarbonate (HCO_3^{-}), chloride (Cl^{-}), and sulphate (SO_4^{2-}). Identification of ions is used in water treatment process to find which are present in water, which undesirable ions to dispose and which ions are needed but not present.

Objectives: Learners will be able to identify cations and anions in a given medium.

a) Identification of cations

Required materials

Apparatus

- Test tubes
- Droppers
- Test tube rack

Chemicals

- Dilute sodium hydroxide solution
- Dilute ammonia solution

Cations: Al^{3+}, Ca^{2+}, Zn^{2+}, Pb^{2+}, Na^{+}, NH_4^{+}, Mg^{2+}, Fe^{2+}, Fe^{3+}, Cu^{2+}

Procedure

1. Put ten test tubes in a test tube rack;
2. Pour one of a solution of each metal ions in test tubes separately (Al^{3+}, Ca^{2+}, Zn^{2+}, Pb^{2+}, Na^{+}, NH_4^{+}, Mg^{2+}, Fe^{2+}, Fe^{3+}, Cu^{2+}) and label them respectively;
3. Divide the solution into two portions;

   To the first portion of each solution, add few drops of NaOH and record your observations;
4. Add more sodium hydroxide until in excess and write down your observations;
5. To the second portion, add few drops of NaOH and record your observations;
6. Add more ammonia solution until in excess and write down your observations;
7. Record all your observations for steps 4-7 in the table below;

Data recording

Cation	Colour of their solution	Addition of aqueous NaOH	Addition of excess aqueous NaOH	Addition of aqueous NH_3	Addition of excess aqueous NH_3
Na^+	Colourless solution	No observable change	No observable change	No observable change	No observable change
Al^{3+}	Colourless solution	White precipitate (ppt)	The ppt dissolves	White precipitate (ppt)	The ppt does not dissolve
Ca^{2+}	Colourless solution	White precipitate	The ppt does not dissolve	No precipitate formed	No precipitate formed
Zn^{2+}	Colourless solution	White precipitate	The ppt dissolves	White precipitate	The ppt dissolves- clear colourless solution
Pb^{2+}	Colourless solution	A white precipitate soluble in excess sodium hydroxide solution gives a colourless solution.	The ppt dissolves to give a colourless solution	White precipitate	The ppt does not dissolve.
NH_4^+	Colourless solution	A gas with a pungent smell that turns red litmus paper blue.	No further reaction	No observable change	No observable change
Mg^{2+}	Colourless solution	White precipitate	The ppt does not dissolve.	White precipitate	The ppt does not dissolve.
Fe^{2+}	Green solution	Dark green precipitate	The ppt does not dissolve	Dark green precipitate	The ppt does not dissolve
Fe^{3+}	Brown solution	Brown precipitate	The ppt does not dissolve	Brown precipitate	The ppt does not dissolve
Cu^{2+}	Blue solution	Blue precipitate	The ppt does not dissolve	Blue precipitate	The ppt dissolves to give a deep blue solution.

Note: Distinguishing between aluminium ion (Al^{3+}) and lead ion (Pb^{2+})

Test: Add potassium iodide (KI) solution

Observations:

For Al^{3+}: No observable change

For Pb^{2+}: Formation of a bright yellow precipitate of PbI_2 [$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$]

Interpretation of results and conclusion

Al^{3+}, Ca^{2+}, Zn^{2+}, Pb^{2+}, Na^+, NH_4^+, Mg^{2+}, Fe^{2+}, Fe^{3+}, Cu^{2+} ions are tested using sodium hydroxide solution. With ammonium ion, a pungent colourless gas is released.

With sodium ion there is no reaction. For other cations, there is formation of the corresponding hydroxides.

Some of them are insoluble in excess are soluble in excess sodium hydroxide whereas others are dissolved. Zinc hydroxide is even dissolved in excess ammonia solution (See table above).

For some cations tests with sodium hydroxide solution or ammonia solution give the same results. Therefore, it is necessary to perform additional specific tests to confirm the presence of these cations in a given medium.

b) Identification of anions

Required materials

Apparatus

- Test tubes
- Droppers
- Bunsen burner

Chemicals

- Aqueous silver nitrate, ($AgNO_3$)
- dilute nitric acid, (HNO_3)
- dilute ammonia (NH_3) solution
- Barium chloride solution
- Conc. sulphuric acid, $H_2SO_4(aq)$
- Dilute hydrochloric acid, HCl(aq)

Anions: Cl^-, I^-, NO_3^-, CO_3^{2-}, SO_3^{2-} and SO_4^{2-}

1) Testing for halides (chlorides, iodides and bromide ions) in solution

Procedure

1. Put three test tubes in a test tube rack;
2. The first test tube contains a solution of Br^-, second a solution of Cl^- and the last a solution of I^-;
3. Put a few drops of dilute nitric acid followed by silver nitrate in each test tube.

   What do you observe?

   Answer: *Test tube which contains Cl^- forms a white precipitate, test tube containing Br^- forms a cream precipitate while one of I^- forms a yellow precipitate.*

2) Test for sulphate (SO_4^{2-}) and sulphite (SO_3^{2-}) ions in solution

Procedure

1. Put two test tubes, one containing sulphate (SO_4^{2-}) and other sulphite (SO_3^{2-}) ions, in a test tube rack;
2. Add barium chloride solution to the test tubes and shake the mixture;

   What do you observe?

   Answer: *A white precipitate is formed in both test tubes.*
3. Add dilute nitric acid to the solution and shake the mixture.

   What do you observe?

   Answer: *The precipitate dissolves in a test tube containing SO_3^{2-} and does not dissolve in the test tube containing SO_4^{2-}.*

3) Testing for nitrates (NO_3^-) in solution

Procedure

1. Put a solution containing nitrate ion in a test tube;
2. Put a few drops of concentrated sulphuric acid in the test tube;
3. Heat the test tube gently on Bunsen burner.

   What do you observe:

   Answer: *Brown fumes are observed.*

4) Testing for carbonates (CO_3^{2-}) ions.

Procedure

1. Put a solution containing CO_3^{2-} ion in a test tube;
2. Add dilute hydrochloric acid to the sample and shake the mixture.

   What do you observe?

   Answer: *There is effervescence in the test tube.*
3. Test the gas produced by bubbling it through lime water.

   What do you observe?

   Answer: *White precipitate is formed (the solution turns milky).*

Interpretation of results and conclusion

Guiding questions:

1. Describe deeply identification tests of halide, sulphate, nitrate and carbonate ions.

When aqueous silver nitrate ($AgNO_3$) followed by excess dilute nitric acid is added to salt solution Cl^- ions form a white precipitate of AgCl, Br^- ions form a pale yellow (cream) precipitate of AgBr, I^- ions form yellow precipitate of AgI. This test indicates that each of these ions is present.

When aqueous barium chloride solution is added to a solution containing SO_3^{2-} or SO_4^{2-}ion, white precipitates of $BaSO_3$(s) or $BaSO_4$(s) respectively. Barium sulphite dissolves in dilute nitric acid but not barium sulphate.

Nitrates form brown fumes of NO_2(g) with concentrated sulphuric acid after heating.

CO_3^{2-} ion gives effervescence of CO_2(g) with dilute hydrochloric and the gas give a white precipitate of $CaCO_3$(s) with lime water.

Unit 7: The Mole concept and gas laws

Experiment 7.1: Determination of the percentage composition of magnesium and oxygen in magnesium oxide

Objective: Learners will be able to determine the percentage composition of magnesium in magnesium oxide by burning magnesium ribbon in air.

Required materials

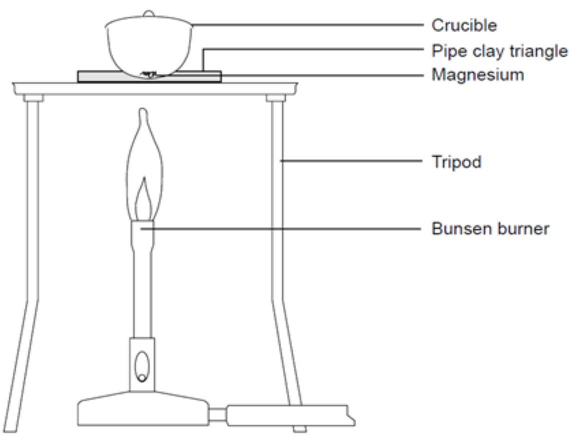
Apparatus

- Safety goggles
- Balance
- Ring stand
- Bunsen burner
- Ring support/ clay triangle
- Crucible/ lid
- Tongs
- Clay tile

Chemicals

- Magnesium ribbon, Mg(s)

Experiment set up



Caution

- Eye protection is essential;
- Do not breathe the fumes generated;
- Do not touch the crucible, lid, triangle, ring, or stand during or after they have been heated;
- Never place anything hot on a balance;
- Do not look into the crucible when it is heating.

Procedure

1. Heat the empty crucible and lid for about 3 minutes to remove water, oils, or other contaminants;
2. Record the mass of crucible;
3. Put about 0.3 g magnesium ribbon in the crucible;
4. Record the mass of the magnesium ribbon and the crucible;
5. Place the crucible securely on the clay triangle. Set the lid slightly off-center on the crucible to allow air to enter but to prevent the magnesium oxide from escaping;
6. Place the Bunsen burner under the crucible, light it and heat strongly until all the magnesium change completely the colour;
7. Stop heating and allow the crucible, lid and contents to cool;
8. Weigh the mass of the crucible and contents. The contents are your new compound of magnesium oxide.

Data recording:

Mass of magnesium ribbon = **X** g

Mass of crucible = **W** g

Mass of crucible and contents after heating = **Z** g

Calculations:

1. The mass of the contents of the crucible after heating is (z-w) g. This is your new compound of magnesium oxide;

2. The percentage of magnesium in the compound can be calculated as:

   $= \frac{X \times 100}{Z\text{-}W}$

3. The mass of oxygen in Magnesium oxide is [(Z-W)-Z g;

4. The percentage of oxygen can be calculated as: $\frac{[(Z\text{-}W)\text{-}X] \times 100}{Z\text{-}W}$

Conclusions

The percentage composition of magnesium oxide is………. of magnesium and ……...of oxygen.

Error Analysis

The above answers are your actual values. If the formula for magnesium oxide is MgO use your periodic table to calculate the theoretical percentage composition of both the magnesium and the oxygen.

The percentage composition for MgO is ………. of magnesium and ………………of oxygen.

% error = $\frac{Theoretical.values - Actual.value...(Experimental)}{Theoretical.value}$

Calculate the % error for O: ……………………………………

Calculate the % error for Mg……………………………………

Interpretation of results and conclusion

1. Write a balanced equation of the reaction magnesium metal and oxygen.

In this experiment, magnesium metal is oxidised by oxygen gas to magnesium oxide. Magnesium reacts vigorously when heated in the presence of air.

$Mg(s) + O_2(g) \rightarrow MgO(s)$

Unit 8: Preparation and classification of oxides

Experiment 8.1: Reaction of oxygen with metals and non-metals

Objective: Learners will be able to prepare oxides by direct combination of oxygen and metals and non-metals.

Required materials

Apparatus

- Pair of tongs
- Deflagrating spoon
- Bunsen burner

Chemicals

- Magnesium ribbon, Mg(s)
- Wood charcoal
- Sulphur powder

a) Reaction with magnesium

Procedure

1. Prepare oxygen gas jar;
2. Wind a piece of a clean Magnesium ribbon on a pencil to form a coil;
3. Using a pair of tongs, ignite the ribbon in Bunsen burner flame;
4. Lower the burning magnesium into the gas jar of oxygen (Ensure the burning Magnesium does not come into contact with the sides of the gas jar);

What do you observe?

Answer: *Grey white powder is observed*

b) Reaction with carbon

Procedure

1. Heat a piece of charcoal in a deflagrating spoon until it grows;
2. Lower the burning charcoal into the gas jar of oxygen.

What do you observe?

Answer: *Grey powder and evolution of a gas is observed*

c) Sulphur

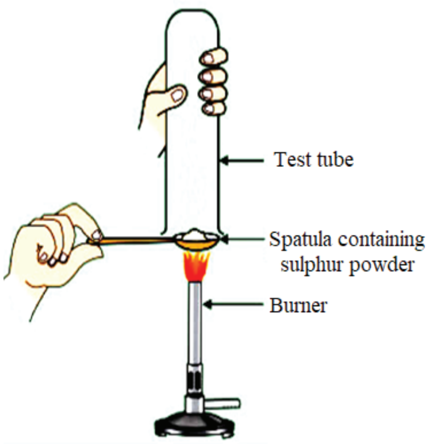
Procedure

1. Place some sulphur powder in a deflagrating spoon;
2. Heat the powder in a deflagrating spoon until it grows;
3. Lower the burning sulphur into the gas jar of oxygen.

What do you observe?

Answer: *Grey powder and evolution of colourless gas is observed.*

Experiment set-up



Interpretation of results and conclusion

Guiding questions

1. Write chemical equations of the reactions when magnesium, sulphur and carbon react with temperature.

Some chemical elements burn in oxygen to produce oxides. Metals react with oxygen to produce metal oxides.

Magnesium reacts with oxygen on heating to form magnesium oxide.

$$2Mg\ (s) + O_2(g) \rightarrow 2MgO\ (s)$$

Similarly, non-metals are also burn in air to produce non-metal oxides. Sulphur combines with oxygen to give sulphur dioxide.

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Carbon combines with oxygen to give carbon dioxide.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Experiment 8.2: Thermal decomposition of hydroxides, carbonates and nitrates.

Objective: Learners will be able to prepare oxides from thermal decomposition of hydroxides, carbonates and nitrates.

Required materials

Apparatus

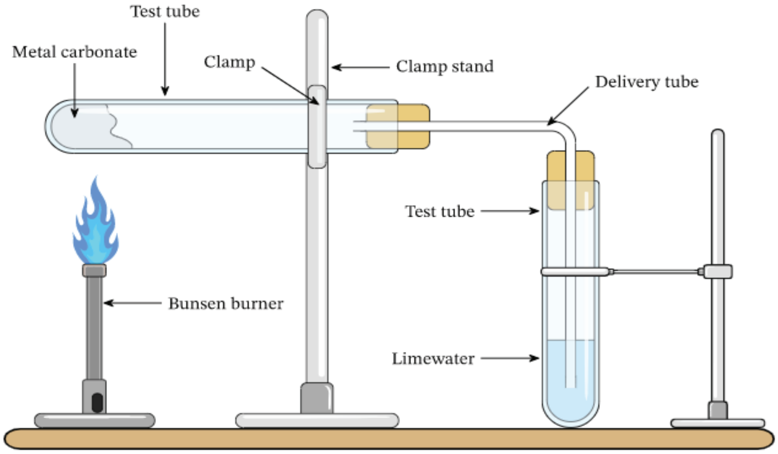
- Boiling test tubes
- Delivery tubes
- Bunsen burner
- Test tube holder

Chemicals

- Copper (II) carbonate (s)
- Lead nitrate (s)
- Copper (II) hydroxide (s)

a) Thermal decomposition of copper carbonate

Experiment set up



Procedure

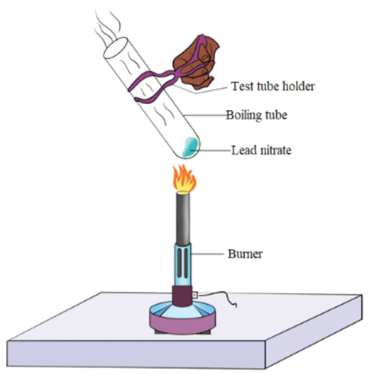
1. Take a little amount of copper carbonate in a boiling test tube;
2. Arrange all apparatuses as shown in the figure above;
3. Heat the test tube on Bunsen burner;

   What do you observe?

   Answer: *The blue-green copper carbonate turns black after heating and a colourless gas is evolved.*

b) Thermal decomposition of lead nitrate

Experiment set up



Caution

- Wear goggles, do not look directly into the boiling test tube;
- Do not touch the boiling test tube;
- Hold the boiling test tube with pair of tongs properly.

Procedure

1. Take about 0.5 g Lead nitrate powder in a boiling test tube;
2. Hold the boiling test tube with a pair of tongs;
3. Heat it over a flame, as shown in figure above;
4. Repeat the previous experiment by using copper hydroxide and record your observation.

   What do you observe?

   Answer: *The blue copper nitrate turns black after heating and brown fumes are observed. In case of copper (II) hydroxide, it turns black after heating and water vapour is observed.*

Interpretation of results and conclusion

Guiding questions

1. Write chemical equations of the thermal decomposition of copper carbonate, copper nitrate and copper hydroxide according to your observations

When copper carbonate (blue-green) is heated, it decomposes into copper oxide (black) and carbon dioxide.

$$CuCO_3(s) \xrightarrow{\text{Heat}} CuO(s) + CO_2(g)$$

Copper nitrate decomposes on gentle heating to give copper oxide, oxygen and brown fumes of nitrogen dioxide are released.

$$2Cu(NO_3)_2(s) \xrightarrow{\text{Heat}} 2CuO(s) + 4NO_2(g) + O_2(g)$$

Copper hydroxide on heating decomposes into copper oxide and water.

$$Cu(OH)_2(s) \xrightarrow{\text{Heat}} CuO(s) + H_2O(g)$$

Experiment 8.3: Properties of oxides (acidic, basic, neutral and amphoteric oxides)

Objective: Learners will be able to identify basic, acidic, neutral and amphoteric oxides.

Required materials

Apparatus

- Boiling test tubes
- Deflagrating spoon
- Bunsen burner
- Gas jar and gas jar lid
- Test tube holder
- Spatula
- Reagent bottles
- Beakers
- Glass rod
- Round bottomed flask
- Match box
- Trough

Chemicals

- Sulphur
- Magnesium ribbon
- Water
- Lead nitrate (Pb $(NO_3)_2(s)$)
- Litmus papers (blue and red)
- Aluminium oxide, $Al_2O_3(s)$
- Hydrochloric acid, $HCl(aq)$
- Sodium hydroxide, $NaOH(aq)$
- Universal indicator
- Ammonium nitrate

Procedure

a) Test for acidic, basic and neutral oxides

1. Ignite a small amount of powdered sulphur on a deflagrating spoon and insert in a gas jar of oxygen;
2. Add 10 mL water after ignition is complete and cover the gas jar with a lid and shake;
3. Take three test tubes and pour 3 mL of the solution to each test tube;
4. Add a few drops of universal indicator solution to the first test tube;
5. Add red and blue litmus paper in the second and third test tubes.

   What do you observe?

   Answer: *The solution in the test tube turns red when universal indicator is added, turns blue litmus paper red and no change when red litmus paper is added.*

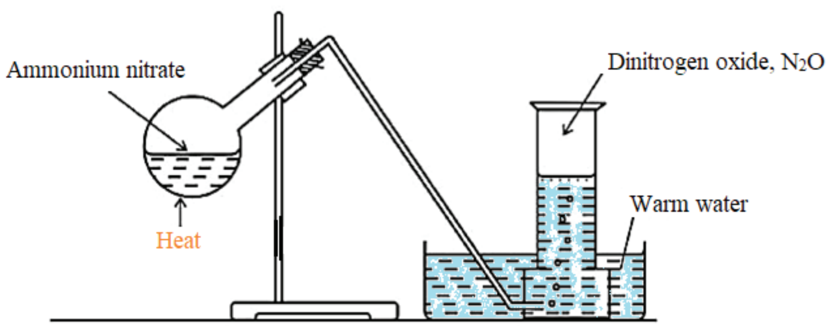
6. Repeat the experiment above by using Magnesium ribbon.

   What do you observe?

   Answer: *The solution obtained after addition of 10mL of water turns blue on addition of universal indicator, turns blue red litmus paper and no change when blue litmus paper is added.*

b) Test for neutral oxides

Experiment set-up



Procedure

1. Take a little amount of ammonium nitrate in a boiling flask;
2. Arrange all apparatuses as shown in the figure above;
3. Heat strongly ammonium nitrate on Bunsen burner until it decomposes.
4. Test the gas evolved by wet blue and red litmus paper. Lightly dampen a red litmus paper test strip with distilled water. Do not use tap water.

   What do you observe?

   Answer: *The evolved gas does not have effect on both blue and red litmus paper.*

c) Test for amphoteric oxides

1. Mix 20 mL of concentrated HCl and 80 mL of water in one reagent bottle;
2. Dissolve 8 g NaOH in 100 mL water in another reagent bottle;
3. Take two beakers and place a spatula full of Al_2O_3 in each beaker;
4. Pour the HCl solution into one beaker and NaOH solution into the other. Stir the mixture with a glass rod;
5. Divide the solutions in 6 test tubes (three for each solution);
6. Add a few drops of universal indicator solution in each two test tubes one of NaOH and HCl;
7. Add red litmus paper in each of the following two test tubes one of NaOH and HCl;
8. Add blue litmus paper in each of the last two test tubes one of NaOH and HCl.

   What do you observe?

Answer:

Product of Al_2O_3 with	Red Litmus paper	Blue litmus paper	Universal indicator
NaOH	Turns blue	Remain unchanged	The solution turns blue/ violet
HCl	Remain unchanged	Turns red	The solution turns red

Interpretation of results and conclusion

Guiding questions

1. What type of oxide formed when:

   a) Non-metal oxides dissolve in water?

   b) Metal oxides dissolve in water?

2. Write equation of reactions in a) and b) above`

3. Write equation of the reaction of the decomposition of ammonium nitrate, the reactions of aluminium oxide react with sodium hydroxide and hydrochloric acid. What type of oxide is aluminium oxide?

A solution of sulphur dioxide and water turns red when universal indicator is added and turns blue litmus paper red, means that it is acidic. Non-metal oxides dissolve in water to form acids. Sulphur dioxide reacts with water to give sulphurous acid meaning that it is acidic oxide.

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

A solution of magnesium oxide and water turns blue on addition of universal indicator and turns red litmus paper blue showing that it is basic. The metal oxides dissolve in water to form base. Magnesium oxide reacts with water to give magnesium hydroxide meaning that it is basic oxide.

$$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$$

When ammonium nitrate is heated, it decomposes to nitrous oxide and water.

Equation of the reaction:

$$NH_4NO_3(s) \xrightarrow{\text{heat, }170^{o}C} N_2O(g) + 2H_2O(g)$$

The gas (N_2O) does not affect blue and red litmus paper. In addition, it has no effect on the universal indicator. The gas produced does not dissolve in water to form a basic or acidic solution. Hence it is neutral compound. Some non-metal oxides do not show any basic or acidic property, they are neutral oxides.

Some metal oxides show both acidic and basic properties. The metal oxides which show both acidic and basic nature are called amphoteric oxides. Amphoteric oxides react with both acids and bases to form salt and water.

Aluminium oxide reacts with hydrochloric acid to form aluminium chloride (salt) and water.

$Al_2O_3(s)$	+	$6HCl(aq)$	$\longrightarrow$	$2AlCl_3(aq)$	+	$3H_2O(l)$
Aluminium oxide		Hydrochloric acic		Alumunium chloride(salt)		Water

In this reaction, aluminium oxide behaves as a basic oxide because it reacts with an acid to form salt and water.

Aluminium oxide reacts also with sodium hydroxide to form sodium aluminate (salt) and water.

$Al_2O_3(s)$	+	$2NaOH(aq)$	$\longrightarrow$	$2NaAlO_2(aq)$	+	$H_2O(l)$
Aluminium oxide		Sodium hydroxide		Sodium aluminate(salt)		Water

In this reaction, aluminium oxide behaves as an acidic oxide because it reacts with a base to form salt and water. Aluminium oxide is amphoteric oxide.

Unit 9: Electrolytes and non-electrolytes

Experiment 9.1: Electrolytes and non-electrolytes

Objective: Learners will be able to distinguish between electrolytes and non-electrolytes.

Required materials

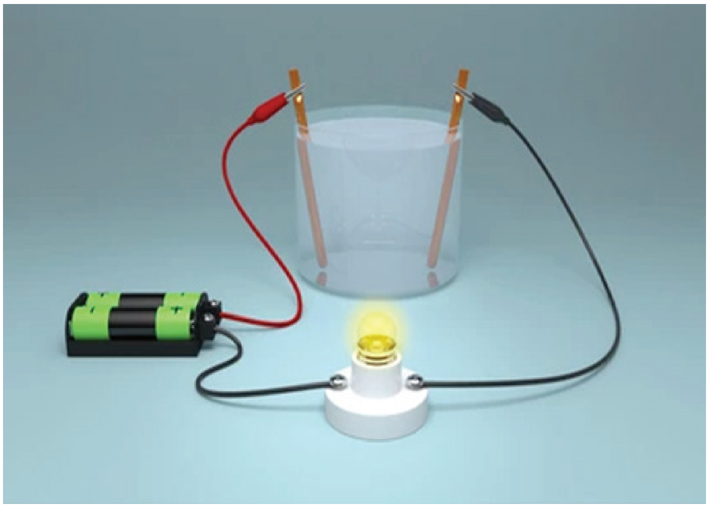
Apparatus

- Batteries
- Conducting wires
- Light bulb
- Graphite electrodes

Chemicals

- Dilute hydrochloric acid
- Sugar solution

Experiment set up



Procedure

1. Take 2 beakers. Label them A and B;
2. Half-fill the beaker A with dilute hydrochloric acid and B with sugar solution;
3. Set all apparatuses as shown in the figure;
4. Introduce the graphite electrodes in the hydrochloric acid solution.

   What do you observe?

   Answer: *The bulb lights up*
5. Wash the electrodes and then dry them with a piece of tissue;
6. Repeat the procedure using the sugar solution.

   What do you observe?

   Answer: *The bulb does not light up.*

Interpretation of results and conclusion

Guiding equations

1. Explain why hydrochloric acid conducts electricity while sugar solution does not conduct.
2. Which is electrolyte and which is a non-electrolyte between hydrochloric acid and sugar?

In the experiment, only hydrochloric acid solution conducts electricity because the solution contains ions, hydrochloric acid ionizes into H^+ and Cl^- in water. It is an electrolyte. Sugar solution does not contain any ion. So, it does not conduct electricity. Sugar is a non-electrolyte.

Experiment 9.2: Classification of substances into strong electrolytes, weak electrolytes and non-electrolytes

Rationale: Electrolyte solutions are ubiquitous/omnipresent. Common examples such as residential tap water and saltwater are used by billions of people every day. There are also less familiar examples that play essential roles in biological, environmental, and industrial systems worldwide. Many of the key examples of these are sodium chloride, potassium chloride, calcium phosphate, magnesium sulphate (Epsom salt), hydrochloric acid, sulphuric acid, acetic acid (vinegar) have many important biological, environmental and industrial uses. Non electrolytes play also important roles in biological, and industrial processes.

Objective: Learners will be able to classify substances into strong electrolytes, weak electrolytes and non-electrolytes.

Required materials

Apparatus

- Batteries
- Conducting wires
- Light bulb
- Graphite electrodes

Chemicals

- Distilled water
- Sodium chloride solution
- Dilute hydrochloric acid
- Sugar solution
- Dilute ethanoic acid
- Dilute ethanol

Experiment set up

Distilled water

Sugar solution

Sodium chloride solution

Procedure

1. Take 6 beakers, and half fill them with the given solution separately;
2. Label the beakers;
3. Introduce the graphite electrodes in each solution respectively. Remember to wash and dry the electrodes before you introduce in the next solution.

   What do you observe?

   Answer:

Solution	What happens to the bulb
Sodium chloride solution	Bulb glows brightly
Dilute hydrochloric acid	Bulb glows brightly
Sugar solution	Bulb does not grow
Dilute ethanoic acid	Bulb glows faintly
Ethanol	Bulb does not grow
Distilled water	Bulb does not grow

Interpretation of results and conclusion

Guiding questions

1. Explain how to show experimentally that sodium chloride and hydrochloric acid are strong electrolytes, acetic acid is a weak electrolyte and ethanol, distilled water and sugar are non electrolytes.

The bulb connected to the beaker containing sodium chloride and hydrochloric acid solutions lights up brightly. This shows that sodium chloride and hydrochloric acid solutions are strong electrolyte, as they ionize completely in solution to produce ions. These ions carry current, and cause the bulb to light up brightly.

The bulb connected to the beaker containing acetic acid solution lights up faintly. This shows that acetic acid is a weak electrolyte, as it ionizes only partially. The solution contains very few numbers of ions. So, the bulb lights up faintly.

The bulb connected to the beaker containing ethanol, distilled water and sugar solution does not light up at all, means that those solutions are not electrolytes; they do not ionize at all to produce ions.

Bulb does not glow | Bulb glows brightly | Bulb glows faintly

Unit 10: Properties of organic compounds and uses of alkanes

Experiment 10.1: Experiment distinguish organic from inorganic compounds in terms of flammability, volatility and in water

Rationale: As safety is an important factor, it's important to consider the flammability of the liquid to be heated. Almost all organic liquids are considered "flammable," meaning they are capable of catching on fire and sustaining combustion (an important exception is that halogenated solvents tend to be non-flammable). However, this doesn't mean that all organic liquids will immediately **ignite** if placed near a heat source. Another important property in discussing flammability is a liquid›s **autoignition temperature:** the temperature where the substance spontaneously ignites under normal pressure and **without** the presence of an ignition source. All liquids with low **autoignition temperature** should be treated more cautiously.

Objective: Learners will be able to differentiate between organic compounds and inorganic compounds.

a) Flammability test

Required materials

Apparatus

- Matchstick/lighter
- Watch glass
- Match box

Chemicals

- Ethanol
- Water

Procedure

1. Place a few drops of ethanol, hexane, ethanol, acetone (organic compounds) and water sodium chloride, calcium carbonate (mineral compounds) in a separate watch glasses;

2. Light a matchstick and burn the substances above on the watch glass.

   What do you observe?

   Answer: *Organic compounds burn, inorganic compounds do not burn.*

Interpretation of results and conclusion

Most inorganic compounds contain ionic bonds - atoms tightly held together in contrast to organic (carbon) compounds. This allows organic compounds to react with oxygen. Salts, inorganic compounds, do not react with oxygen, hence they are non-combustable.

Generally, organic compounds are flammable, they burn in oxygen. Inorganic compounds are not flammable.

b) Solubility test

Required materials

Apparatus

- Test tube
- Beaker
- Dropper

Chemicals

- Water
- Chloroform
- Naphthalene
- Sodium chloride

Procedure

1. Put chloroform in 2 separate test tubes;
2. Add a pinch of naphthalene in the first test tube and sodium chloride in the second test tube.

   What do you observe?

   Answer: *Naphthalene dissolves in chloroform but sodium chloride does not.*
3. Place a distilled water in 2 separate test tubes;

4. Add a pinch of naphthalene in a one test tube and sodium chloride in another test tube.

   What do you observe?

   Answer: *White precipitate dissolves when NaCl is added and nothing happens when naphthalene is added.*

Interpretation of results and conclusion

Naphthalene is not soluble in water but it is soluble in chloroform.

Sodium chloride is soluble in water but not soluble in chloroform.

Generally, organic compounds are soluble in organic solvents while inorganic compounds are soluble in inorganic solvents.

b) Volatility

Required materials

Apparatus

- Crucible
- Bunsen Burner
- Spatula

Chemicals

- Ethanol
- Sodium hydroxide, NaOH (s)

Procedure

1. Put 0.5 of ethanol into crucible and 0.3ml of sodium hydroxide wait for 2 min

   What do you observe?

   Answer: *After a while ethanol will be evaporated which indicates that it's a volatile substance whereas sodium hydroxide does not evaporate.*

Interpretation of results and conclusion

	Organic compounds	Inorganic Compounds
1	Generally, they are insoluble in water but soluble in organic solvents.	They are generally soluble in water and non-soluble in organic solvents.
2	They are generally highly inflammable.	They are non-inflammable.
3	They are volatile in general	They are not volatile

Experiment 10.2: Laboratory preparation of methane gas

Objective: Learners will be able to prepare and collect methane gas.

Required materials

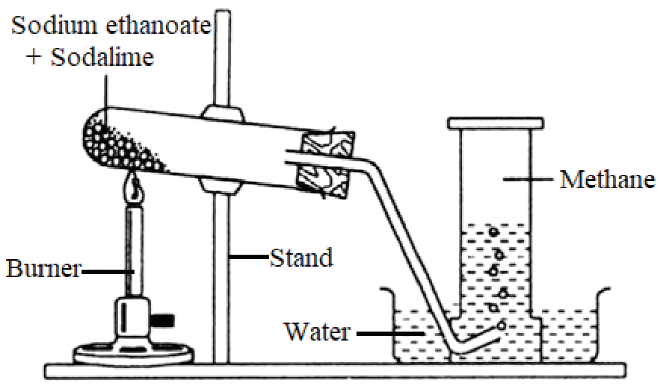
Apparatus

- Stand and accessories
- Delivery tube
- Gas jar
- Boiling tube
- Stopper with one hole
- Bunsen burner
- Electronic balance

Chemicals

- Sodium hydroxide, NaOH (s)
- Sodium acetate, CH_3COONa (s)
- Calcium oxide
- Water

Experiment set up



Procedure

1. Mix sodium hydroxide (NaOH) and calcium oxide (CaO) in the ratio 3:1 to make soda-lime;
2. Take a mixture of anhydrous sodium ethanoate and soda-lime in the ratio 2:1 and transfer the mixture in the boiling tube;
3. Close the boiling tube with a stopper with a gas-delivery tube;
4. Fix the apparatus as shown by the experiment setup above;
5. Heat the boiling tube gently with the cold part of the flame to avoid local overheating and keep the flame in motion.

   What do you observe?

   Answer: *After a while the colourless gas starts liberating.*

   Why the gas is collected downward?

   Answer: *The gas is insoluble and less dense than water.*

Interpretation of results and conclusion

Guiding question:

1. Write a balanced equation of the preparation f methane and show all the physical states of reactants and products.

When the mixture of sodium acetate and soda lime is heated, there is a production of a methane gas and sodium carbonate.

$$\underset{\text{Sodium ethanoate}}{CH_3COONa(s)} + \underset{\text{Sodium hydroxide}}{NaOH(s)} \xrightarrow{\text{Heat}} \underset{\text{Methane}}{CH_4(g)} + \underset{\text{Sodium carbonate}}{Na_2CO_3(s)}$$

CHEMISTRY EXPERIMENTS FOR SENIOR THREE

Experiment 1.1: Preparation of calcium carbonate and lead carbonate

Rationale: This experiment will develop experimental ability of learners and will increase their familiarity with laboratory practices regarding the preparation of carbonates salts. Plants grow better at their specific pH. The overuse of soil may lead to decreased pH hence the soil may become acidic. This experiment aims to produce carbonates which act as pH regulators in order to enhance the soil productivity.

Objective: Learners will be able to prepare calcium carbonate and lead carbonate.

Required materials

Apparatus

- Test tubes
- Spatulas
- Erlenmeyer
- Filter funnel
- Filter paper
- Conical flask
- Beakers
- Stirring rod

Chemicals

- Calcium chloride ($CaCl_2$)
- Sodium carbonate (Na_2CO_3)
- Water
- Lead nitrate ($Pb(NO_3)_2$)

Procedure:

1. Place 5 g of calcium chloride in a beaker and add to it 20 cm^3 of water.
2. Stir until all the solid dissolves.
3. Put 5 g of sodium carbonate in another beaker and add to it 20 cm^3 of distilled water.
4. Stir the mixture to form a uniform solution.
5. Add the sodium carbonate solution to the calcium chloride solution.
6. Stir and allow the mixture to settle.

   What do you observe?

   Answer: *A white precipitate is formed.*
7. Filter the precipitate obtained and wash it with distilled water.
8. Dry the precipitate between filter papers.
9. Repeat steps 1 to 5 using lead (II) nitrate instead of calcium chloride.

   What do you observe?

   Answer: *A white precipitate is formed.*

Interpretation of results

Guiding questions

1. When calcium chloride reacts with sodium carbonate a white precipitate of calcium carbonate is formed. Write equation of this reaction.
2. When lead nitrate reacts with sodium carbonate a white precipitate of lead carbonate is formed. Write equation of this reaction.

When calcium chloride reacts with sodium carbonate a white precipitate of calcium carbonate is formed.

Equation for the reaction:

$$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaCl\ (aq)$$

When lead nitrate reacts with sodium carbonate a white precipitate of lead carbonate is formed. Equation for the reaction:

$$Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2NaNO_3\ (aq)$$

Experiment 1.2: Preparation and test for carbon dioxide

Objective: learners will be able to prepare and test carbon dioxide gas

Required materials

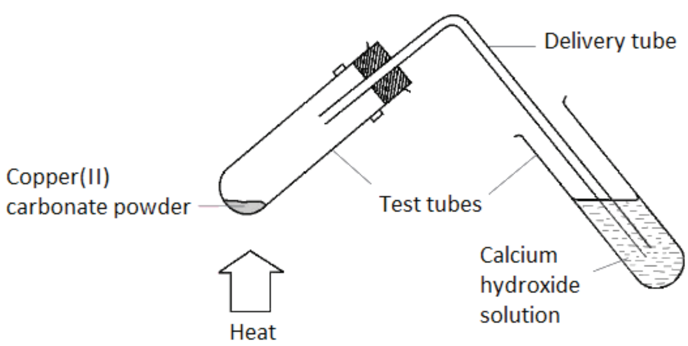
Apparatus

- Test tubes
- Test tube holders
- burning flame

Chemicals

- lime water (calcium hydroxide solution)
- copper carbonate powder

Experiment set up



Procedure

1. Put about 2 spatulas of copper (II) carbonate powder in a dry test tube,
2. Put calcium hydroxide solution half-way the boiling test tube,
3. Arrange the set up as shown in the figure 1,
4. Heat the copper (II)carbonate strongly in the test tube.

What happens in the test tubes containing copper carbonate and lime water?

Answer: *Green powder turns into Black as it decomposes forming a colourless gas which turns lime water milky.*

Interpretation of Results and conclusion

Guiding questions

1. How can you test for the presence f carbon dioxide? Which reactant do you use? What is observed?
2. Write equation of the reaction to prepare carbon dioxide form copper carbonate
3. Write equation for the test of carbon dioxide.

We use limewater to test for the presence of carbon dioxide. Limewater is calcium hydroxide dissolved in water. To test for carbon dioxide, we pump the gas that may contain carbon dioxide through the limewater. If carbon dioxide is present in the gas, the limewater will turn cloudy (milky). If the gas passes for a long time the milky solution turns to colourless which is calcium hydrogen carbonate solution

Equations

$$CuCO_3(s) \xrightarrow[Heat]{} CuO(s) + CO_2(g)$$

Green powder black solid

$$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$$

Cloudy solution/ white ppt

$$CaCO_3(s) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$$

Colourless solution

UNIT 2 Nitrogen and its Inorganic Compounds

Experiment 2.1: Laboratory preparation and test for ammonia gas

Rationale: About 80% of the ammonia produced by industry is used in agriculture as fertilizer. It can be prepared from calcium hydroxide and ammonium chloride while at industry level ammonia is prepared from ammonia of atmosphere and hydrogen.

Objective: Learners will be able to prepare and test ammonia gas by reacting calcium hydroxide and ammonium chloride

Required materials

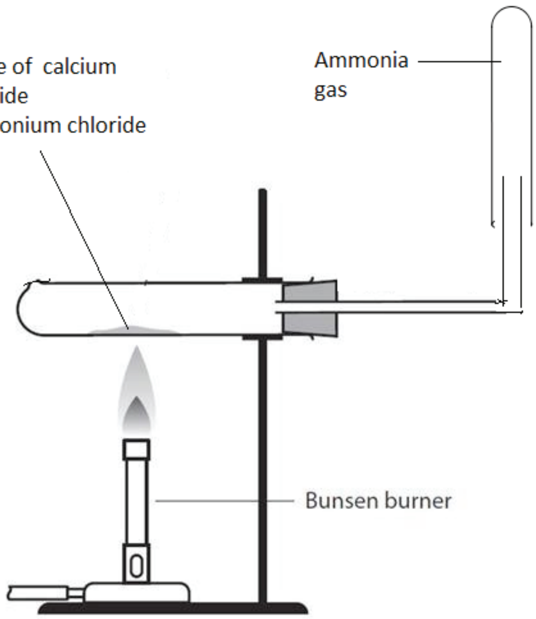
Apparatus

- Mortar and pestle
- Spatula
- Glass rod
- Bunsen burner
- Gas jars
- Retort stand and accessories
- Delivery tubes
- Pyrex test tube for heating

Chemicals

- Ammonium chloride
- Calcium hydroxide
- Distilled water

Experiment set up



Procedure

1. Mix ammonium chloride and calcium hydroxide in the ratio of 2:1.
2. Grind mixture in a mortar and add a little water and mix to get a paste.

   Why is it necessary to add water?

   Answer: *The addition of water quickens the reaction.*
3. Arrange the apparatus as show in figure above
4. Heat the mixture in the flask gently.

   What do you observe?

   Answer: *When the mixture is heated, water droplets containing ammonia gas are seen on the cooler parts of the test tube.*

Tests for ammonia gas:

1. Smell the gas by wafting it towards your noise.
2. Hold a piece of red litmus paper on the mouth of the test tube containing ammonia?

What do you observe?

Answer: *A pungent colorless gas is released, and the red litmus paper turns blue in presence of this gas.*

3. Hold another test tube containing hydrochloric acid near the test tube containing ammonia gas.

   What do you observe?

   Answer: *white fumes of ammonium chloride form near the mouth of ammonia gas*

Interpretation of results and conclusion

Guiding equations

1. Write equation of the reaction showing the preparation of ammonia from calcium hydroxide and ammonium chloride
2. Describe identification test of ammonia using red litmus paper and hydrogen chloride

Calcium hydroxide reacts with ammonium chloride to form calcium chloride, water, and ammonia gas. The equation for the reaction:

$$Ca(OH)_2(aq) + 2NH_4Cl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l) + 2NH_3(g)$$

When a damp red litmus paper is placed into the test-tube of ammonia, it turns blue. This shows that ammonia is basic/alkaline

Ammonia reacts with hydrogen chloride gas to form ammonium chloride

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

Evaluation activity

1. Why should the flask be in a sloping position?
2. Name the method used to collect ammonia gas and suggest why it is preferred.

Experiment 2.2: Demonstration of solubility of ammonia gas in water and pH changes

Objective: learners will be able to demonstrate the solubility of ammonia gas in water.

Required materials

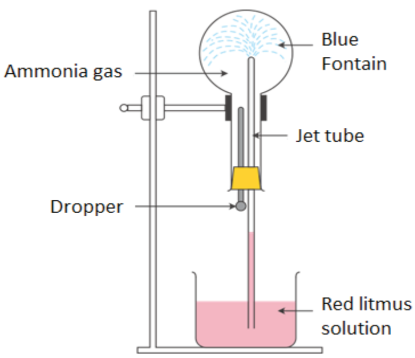
Apparatus

- Round bottomed flask
- Dropper
- Retort flask and its accessories
- Beaker
- Jet tube or syringe

Chemicals

- Red litmus solution
- Ammonia

Experiment set up: Fountain experiment to demonstrate solubility of ammonia



Procedure:

1. Fill the beaker with red litmus solution
2. The top round bottomed flask is filled with ammonia.
3. Arrange the apparatus as shown in figure 3.
4. Squeeze the dropper to transfer some drops of water into the upper flask.

   What do you observe?

   Answer: *The water rushes up to the top flask and the solution turns pink.*

Interpretation and conclusion

The first few drops of water that enter the upper flask absorb a little ammonia gas. When this occurs, a slight vacuum that forms suck a little more water into the flask to balance the pressure. This water, in turn, absorbs a little more ammonia. This continuous exchange creates a chain reaction that causes the solution to come rushing into the top flask.

Experiment 2.3: Laboratory preparation of nitrogen

Rationale: Nitrogen gas is used in laboratory or hospital to create inert atmosphere. Nitrogen is also useful for preparing other nitrogenous compounds such as ammonia, nitric acid that are useful for manufacturing of fertilizers. The modern agriculture needs high quantity of fertilizer, nitrogen is used as nitrogen-rich fertilizer and as such is of great importance in agriculture.

Objective: Learners will be able to prepare and describe the preparation of nitrogen gas.

Required materials

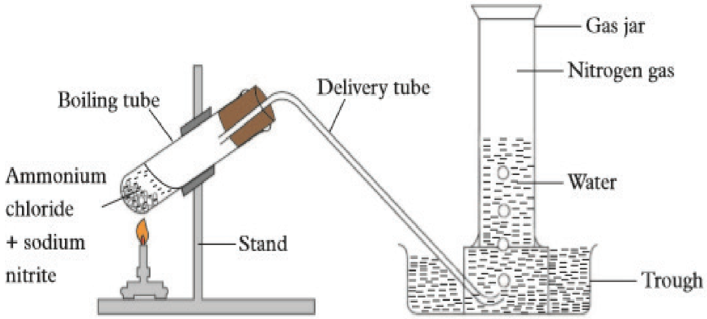
Apparatus

- Round-bottomed flask or boiling tube
- Bunsen burner/any source of heat
- Trough
- Gas jars,
- Delivery tubes
- Retort stand
- Glowing splint

Chemicals

- Sodium nitrite,
- Ammonium chloride

Experiment set up



Procedure:

1. Make a solution of sodium nitrite and ammonium chloride in water
2. Pour the solution in a round-bottomed flask or boiling tube,
3. Arrange the apparatus as shown in the figure below,
4. Heat the flask slightly and then remove the source of heat.

   Record your observations in the flask and gas jar.

   Observation: A colourless gas insoluble in water is evolved

Interpretation of results

Guiding questions

1. What reaction that occurs in the above experiment?
2. Write equation of this reaction and overall reaction.

Sodium nitrite **($NaNO_2$)** reacts with ammonium chloride to form ammonium nitrite and sodium chloride.

$$NaNO_2\ (aq) + NH_4Cl\ (aq) \rightarrow NaCl\ (aq) + NH_4NO_2\ (aq)$$

The ammonium nitrite is then decomposed to produce nitrogen gas.

$$NH_4NO_2\ (aq) \rightarrow N_2(g) + 2H_2O(l)$$

The overall equation for the reaction:

$$NaNO_2\ (aq) + NH_4Cl\ (aq) \rightarrow NaCl(aq) + 2H_2O(l) + N_2(g)$$

Experiment 2.4: Laboratory Preparation of Nitrogen Dioxide

Rationale: Nitrogen dioxide is useful for preparing fertilizers. The fertilisers that are made from nitrogen dioxide increases the production of crops in agriculture.

Objective: Learners will be able to prepare and explain the preparation of nitrogen dioxide.

Required materials

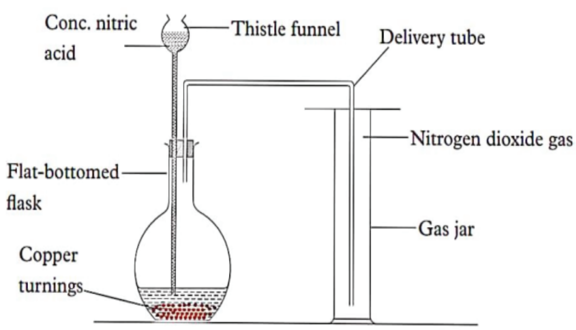
Apparatus

- Thistle funnel or dropping funnel
- Flat-bottomed flask
- Delivery tube
- Gas jars
- Cardboard cover

Chemicals

- Copper turnings
- Concentrated nitric acid.

Experiment set up



Procedure:

1. Put copper turnings into a flat-bottomed flask.
2. Add concentrated nitric acid to the copper turnings in the flask.
3. Open the tap water to cool down the round bottomed flask.
4. Record your observations in the flask and the gas jar.

Observation: After adding concentrated nitric acid to copper turnings a red brown gas is produced and a green solution is also formed in the flask.

Interpretation of the results

Guiding questions

1. What is the brown gas produced and what is the green solution produced?
2. Write equation of the reaction that occurs when copper reacts with concentrated nitric acid to form a green solution and a red-brown gas.

Copper reacts with concentrated nitric acid to produce red brown (nitrogen dioxide) gas.

When concentrated nitric acid is added to copper turnings, effervescence occurs immediately producing red-brown fumes. A green solution is also formed in the flask.

Copper reacts with concentrated nitric acid to form a green solution of copper(II) nitrate, water and a red-brown gas (nitrogen dioxide).

Copper + nitric acid → copper (II) nitrate +water + nitrogen dioxide

$$Cu(s) + 4HNO_3(l) \rightarrow Cu(NO_3)_2(aq) + 2H_2O\ (l) + 2NO_2(g)$$

UNIT 3 Sulphur and its Inorganic Compounds

Experiment 3.1: Laboratory preparation and test for sulphur dioxide

Rationale: Nature sources of sulphur dioxide are volcanoes and fires among others but it can be is easily prepared in laboratory. This gas, SO_2 is widely used in the food and drinks industries for its properties as a preservative and antioxidant. Whilst harmless to healthy persons when used in recommended concentrations, it can induce asthma when inhaled or ingested by sensitive subjects, even in high dilution.

Objective: Learners will be able to prepare sulphur dioxide gas.

Required materials

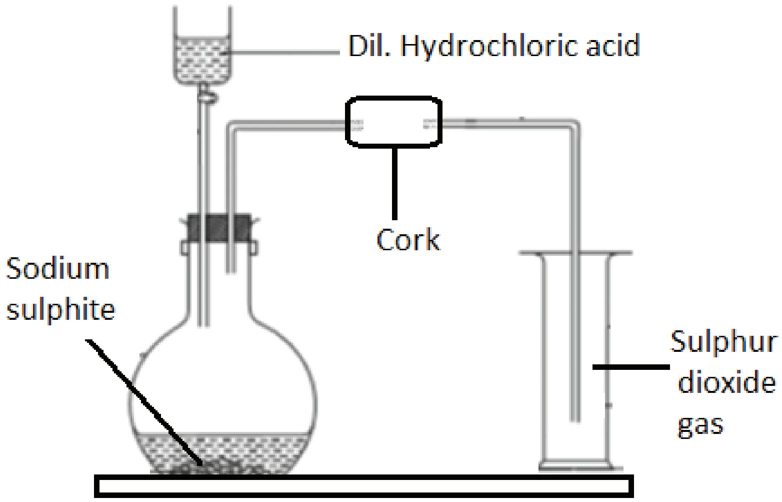
Apparatus

- Flat bottomed flask
- Dropping funnel
- Gas jar
- Wash bottle
- Bunsen burner
- Beakers Distilled water
- Match box

Chemicals

- Sodium sulphite, Na_2SO_3
- Dilute hydrochloric acid, HCl
- Concentrated sulphuric acid, H_2SO_4
- Sodium hydroxide solution, NaOH
- Blue and red litmus papers

Experiment set up



Caution: Sulphur dioxide is poisonous. This experiment should be carried out in a fume chamber.

Procedure:

1. Place sodium sulphite in a flat-bottomed flask and set up the apparatus as shown in figure in the set up above.
2. Run down the hydrochloric acid into the flat-bottomed flask.

   What do you observe?

   Answer: *A colourless gas (sulphur dioxide) flows from the flask.*
3. Collect some sulphur dioxide gas and test it with wet blue and red litmus papers.

   What do you observe?

   Answer: *Blue litmus paper turns red and red litmus paper doesn't change.*

Test of Sulphur dioxide gas

Replace the gas jar successively with test tubes containing about 5 mL of acidified potass um manganate (VII) solution ($KMnO_4$), 5mL of acidified potassium dichromate (VI) solution ($K_2Cr_2O_7$) and 5mL of iron (III) chloride solution ($FeCl_3$).

What do you observe?

Answer: *The solution of potassium permanganate changes from violet to colorless, The solution of potassium dichromate changes from orange to dark green color and solution of iron (III) chloride changes from brown to green.*

Interpretation of results and conclusion

Guiding questions

1. Write equation of the reaction of the preparation of sulphur dioxide.
2. Why sulphur dioxide turns red wet blue litmus paper?

When hydrochloric acid reacts with sodium sulphite, sodium chloride and sulphur dioxide are formed. The sulphur dioxide gas is passed through concentrated sulphuric acid to be dried. It is then collected by downward delivery since it is denser than air.

The equation of reaction:

$$Na_2SO_3\,(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + SO_2(g) + H_2O(l)$$

Sulphur dioxide turns red wet blue litmus paper. This shows that sulphur dioxide is of acidic nature.

Guide for evaluation

1. Explain why heat is used in the beginning?
2. Suggest and discuss another way of preparing sulphur dioxide in laboratory and write down the equation of the reaction.

Experiment 3.2: The dehydration of sugar, pieces of fabric, paper, and wood by concentrated sulphuric acid

Objective: Learners will be able to show the dehydrating properties of sulphuric acid.

Required materials

Apparatus

- Paper towels
- Safety gloves
- 500mL borosilicate beaker
- Spatula Piece of cloth
- Stirring rod
- 250mL graduated cylinder

Chemicals

- 70mL of concentrated sulphuric acid (18M)
- 50g granulated white or brown sugar
- Piece of paper

Experiment set up

Sugar + Sulphuric acid	Colour changes as the reaction occurs	Expansion of Black mass of carbonn

Caution: Concentrated sulphuric acid is very corrosive, therefore hand it carefully.

Procedure

1. Wear safety gloves.
2. Cover the laboratory working table by some paper towels.
3. Put 50g of white sugar into a 500mL beaker.

4. Insert a stirring rod into center of white sugar.
5. Put the beaker on paper towels.
6. Using a graduated cylinder, add 70mL of sulphuric acid to the sugar and stir for a while.
7. Stand in about 1 to 2 meters away and wait for reaction to begin and the column to grow.

   What do you observe?

   Answer:

   - *In about 30 seconds after the concentrated acid is added to the sugar, the color changes (from light, brown and finally to a black mass of substance) as the reaction proceeds.*
   - *The mixture expands as water is removed from the sugar and the beaker becomes very hot.*
   - *A long column of black spongy mass grows up from the beaker.*
   - *There are vapors and the smell of burned sugar.*

8. Pour concentrated sulphuric acid onto a piece of cloth, piece of paper and piece of wood.

   What do you observe?

   Answer:

   - *When concentrated sulphuric acid is poured on piece of paper and piece of cloth, they get burned and torn into small pieces.*
   - *When acid is poured onto piece of wood, black mass is formed.*

Interpretation and conclusion

Guiding equations

1. What happens when concentrated sulphuric acid is added to the sugar.
2. What is responsible of the expansion of the mixture inside the beaker?
3. What is the black mass formed?

When concentrated sulphuric acid is added to the sugar, sugar is stripped of water and leaving behind black carbon (in form of long black spongy substance).

The reaction is highly exothermic and produces water vapors and carbon dioxide. The vaporized water and carbon dioxide are responsible of the expansion of the mixture inside the beaker.

Equation of reaction:

$$C_{12}H_{22}O_{11} \xrightarrow{\text{Conc H2SO}_4} 12C(s) + 11H_2O.H_2SO_4$$

Similarly concentrated sulphuric acid removes water from wood and other organic substances converting them into a black mass which is carbon. These show that concentrated sulphuric acid is a **dehydrating agent.**

Experiment 3.3: Dehydration of hydrated Copper (II) sulphate by concentrated sulphuric acid.

Rationale: Copper sulphate is employed at a limited level in organic synthesis. In addition Copper sulphate is used as a fungicide, algaecide, root killer, and herbicide in both agriculture and non-agricultural settings. It is also used as an antimicrobial.

Objective: Learners will be able to explain the dehydration of copper (II) sulphate by concentrated sulphuric acid.

Required materials

Apparatus

- Test tubes, Concentrated
- Test tube holders,
- Beakers,
- Measuring cylinder,
- Droppers,
- Pair of tongs,
- Spatula.

Chemicals

- Sulphuric acid,
- Copper (II) sulphate crystals.

Procedure:

1. Put 3 spatula endful of blue hydrated copper (II) sulphate crystals in a 50 mL glass beaker and add 5 mL of concentrated sulphuric acid.

   Note down your observations.

   Answer: *the blue crystals of copper (II) sulphate become white.*

Interpretation and conclusion

Guiding question

1. Write a balanced equation of the dehydration of hydrated copper sulphate by concentrated sulphuric acid.

Concentrated sulphuric acid removes water from hydrated copper (II) sulphate (blue) converting it to a white powder of anhydrous copper (II) sulphate.

$$CuSO_4 . 5H_2O_{(s)} \quad CuSO_{4(s)} + 5H_2O_{(l)}$$

Hydrated copper (II) sulphate is dehydrated by concentrated sulphuric acid

Concentrated sulphuric acid is a **dehydrating agent.**

Experiment 3.4: Reaction of dilute sulphuric acid with metals

Rationale: When sulphuric acid reacts with metals, metal sulphates are formed. Many metal sulphates have many important uses: Magnesium sulphate is a naturally occurring mineral used to control low blood levels of magnesium. Magnesium sulphate injection is also used for pediatric acute nephritis and to prevent toxemia of pregnancy.

Calcium sulphate, $CaSO_4$, is a naturally occurring calcium salt. It is employed as a soil conditioner as it helps farmers to improve their soil structure. This type of fertilizer contains all the nutrients required for your plants' growth.

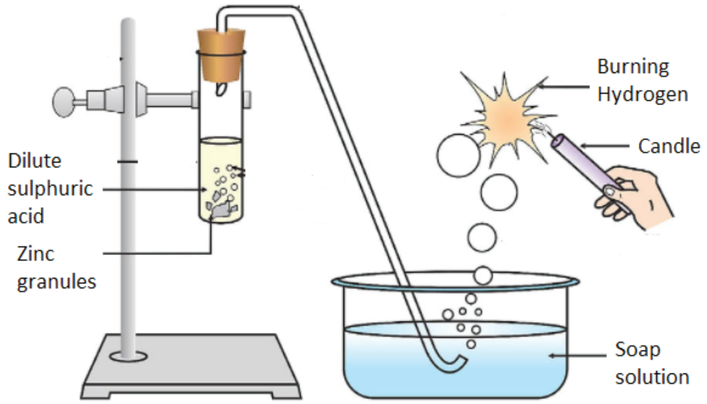
Sodium sulphate is mainly used for **the manufacture of detergents process and** it has many other uses.

Objective: Learners will be able to show and illustrate the reaction of dilute sulphuric acid with metals.

Required materials

Apparatus	**Chemicals**
– Test tubes	– Dilute sulphuric acid
– Wooden splint	– Magnesium ribbon
– Bunsen burner	– Zinc granules
– Match box	– Aluminium

Experiment set up



Procedure

1. Put about 5g of zinc granules in a test tube and add about 3ml of diluted sulphuric acid

   What do you observe?

   Answer: *a colorless gas is evolved*

2. Test the gas evolved using a burning splint.

   Note down your observations.

   Answer: *evolved gas burns with a pop sound*

   What is this gas?

   Answer: *this gas is hydrogen*

3. Repeat the experiment using copper, Magnesium ribbon or powder, and Aluminium instead of Zinc

Interpretation of results and conclusion

Guiding questions

1. In general what happens when diluted sulphuric acid reacts with metals?
2. Compare the reactivity of diluted sulphuric acid reacts with magnesium, zinc, aluminium and copper
3. Write equation of the reaction between diluted sulphuric acid reacts with magnesium

Sulphuric acid reacts vigorously with magnesium and a gas that burns with a pop sound is produced.

Equation of reaction:

$$H_2SO_{4(aq)} + Mg(s) \rightarrow MgSO_{4(aq)} + H_2(g)$$

Sulphuric acid reacts moderately with zinc.

Sulphuric acid reacts slowly with aluminium and small bubbles of hydrogen are released

No gas evolved with copper. Therefore, copper does not react with dilute sulphuric acid.

In general, when diluted sulphuric acid reacts with a metal, it forms metal salt and hydrogen gas.

Equation of reaction:

Metal + Dilute acid $\rightarrow$ Metal salt + Hydrogen gas

Evaluation activity

Write balanced equation of reaction of diluted sulphuric acid with zinc and aluminium

Experiment 3.5: Reaction of concentrated sulphuric acid with metals

Objective: Learners will be able to show the reaction of concentrated sulphuric acid with metals

Required materials

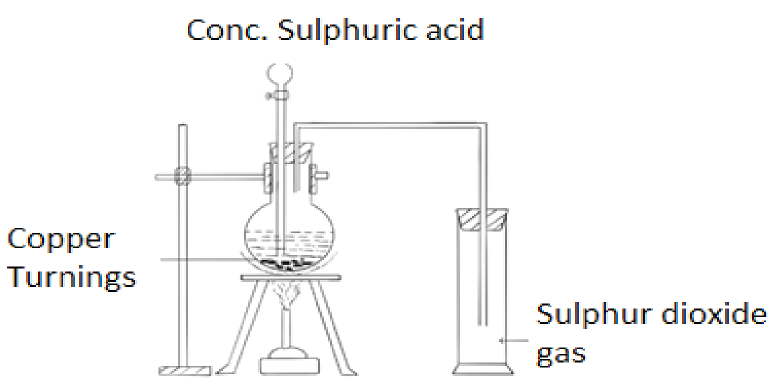
Apparatus

- Round bottomed flask
- Beaker
- Bunsen burner
- Tripod stand
- Wire gauze
- Stand and clamp
- Dropping funnel
- Gas jar
- Delivery tube
- Cork with two holes
- Cork with one hole

Chemicals

- Concentrated sulphuric acid
- Magnesium
- Copper turnings
- Potassium dichromate (VI)

Experiment set up



Procedure

1. Place copper turnings in a test round bottomed flask and add 3mL of concentrated sulphuric acid.
2. Heat the flask gently and test any gas evolved by placing filter paper soaked in acidified potassium dichromate (VI) at the mouth of the gas jar.

   Note down your observations.

   Answer: *Reddish brown copper turnings disappear in the acid forming a blue solution.*

   A colorless shocking gas which turns green the orange potassium dichromate (VI) is released.

Interpretation of results and conclusion

Guiding questions

1. What happens when concentrated sulphuric acid reacts with copper metal? Write corresponding equation.

When concentrated sulphuric acid reacts with copper metal there is formation of copper sulphate, sulphur dioxide (with a shocking smell) and water. The equation of reaction is

$$\underset{\textbf{(Brown)}}{Cu(s)} + 2H_2SO_4(l) \rightarrow \underset{\textbf{(Blue)}}{CuSO_4(aq)} + SO_2(g) + 2H_2O(l)$$

Experiment 3.6: Test of the presence of sulphites and sulphates in solution

Objective: Learners will be able to identify sulphate and sulphite ions in a given solution.

Required materials

Apparatus

- Test tubes,
- Droppers.

Chemicals

- Sodium sulphite,
- Sodium sulphate,
- Barium nitrate,
- Dilute nitric acid,
- Dilute hydrochloric acid.

Procedure

1. Place 2 mL of sodium sulphate solution and 2 mL of sodium sulphite in separate test tubes.
2. Add four drops of barium nitrate solution in each solution in the test tubes

   What do you observe?

   Answer: *a white precipitate is formed.*
3. Add 4 drops of dilute nitric acid to each test tube in procedure 2 and shake,
4. Record the observations in the following table.

Test record

Ion solution	Addition of barium nitrate solution	Addition of dilute nitric acid solution
SO_3^{-2}	White precipitate is formed	Precipitate is dissolves
SO_4^{-2}	White precipitate is formed	Precipitate persists

Interpretation and conclusion:

1. Each of SO_3^{-2} and SO_4^{-2} precipitates in the presence of $Ba^{2+}(aq)$. Write equations of corresponding reactions.

Each of $SO_3^{-2}$and SO_4^{-2} precipitates in the presence of $Ba^{2+}(aq)$.

Equation of the reactions:

$$SO_3^{-2}(aq) + Ba^{2+}(aq) \rightarrow BaSO_3(s)$$

$$SO_4^{-2} (aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$$

In the presence of dilute nitric acid, BaSO3(s) dissolves while BaSO4(s) does not dissolve.

Note: You can also use barium chloride solution and Hydrochloric acid respectively instead barium nitrate and nitric acid.

UNIT 4 Chlorine and its Inorganic Compounds

Experiment 4.1: Preparation and test of chlorine.

Rationale: Chlorine is prepared in laboratory by the reaction between manganese oxide and concentrated hydrochloric acid. This chlorine has a variety of uses. It is used to disinfect water and is part of the sanitation process for sewage and industrial waste. During the production of paper and cloth, chlorine is used as a bleaching agent. It is also used in cleaning products, including household bleach which is chlorine dissolved in water.

Objective: Learners will be able to prepare and test chlorine gas.

Required materials

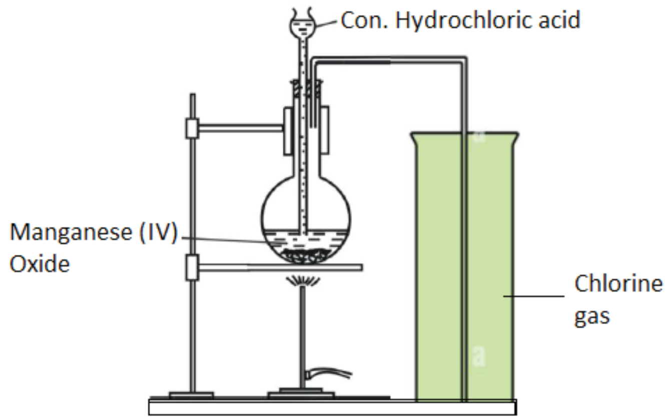
Apparatus

- Round bottom flask
- Conical flask
- Gas jars
- Source of heat
- Wash bottles
- Delivery tubes Water
- Thistle funnel
- Bunsen burner
- Retort stand and clamp
- Match box

Chemicals

- Manganese oxide
- Potassium manganate (VII)
- Concentrated hydrochloric acid
- Blue litmus paper
- Red litmus paper

Experiment set up



Caution: Chlorine is toxic, avoid inhaling it.

Procedure:

1. Put manganese dioxide in a round bottomed flask and set up the apparatus as shown above.
2. Add concentrated hydrochloric acid into the flask using thistle funnel and heat the mixture.

   Note down your observations.

   Answer: *there is effervescence and yellowish vapors are released.*

   What is the role of water and concentrated sulphuric acid in the wash bottles as they are shown in the set up above?

   Answer: *water absorbs hydrogen chloride fumes and concentrated sulphuric acid helps to dry chlorine gas produced. Sulphuric acid is a dehydrating agent.*
3. Test the gas collected with moist blue litmus paper and red litmus paper.

   Note down your observations.

   Answer: *the gas turns red the blue litmus and later bleaches it. Red litmus doesn't change the color but after a while it is bleached.*
4. Smell the gas by wafting it towards your nose.

   Describe the smell of the gas.

   Answer: *the gas can be recognized by its pungent and irritating odor.*

Interpretation of results and conclusion

Guiding question:

1. Describe the preparation of and test of chlorine and write equation of the reaction that occurs during its preparation

Hydrochloric acid reacts with manganese (IV) oxide to produce manganese (II) chloride, Chlorine gas and water. The equation of reaction is:

$$MnO_2(s) + 4HCl \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

The water solution of chlorine is acidic and a bleaching agent.

Experiment 4.2: Laboratory preparation and test of hydrogen chloride

Rationale: Chlorine is prepared in laboratory by the reaction between sodium chloride and concentrated sulphuric acid. This hydrogen chloride has many uses, including cleaning, pickling, electroplating metals, tanning leather, and refining and producing a wide variety of products.

Objective: Learners will be able to prepare hydrogen chloride from concentrated sulphuric acid and sodium chloride and test it.

Required materials

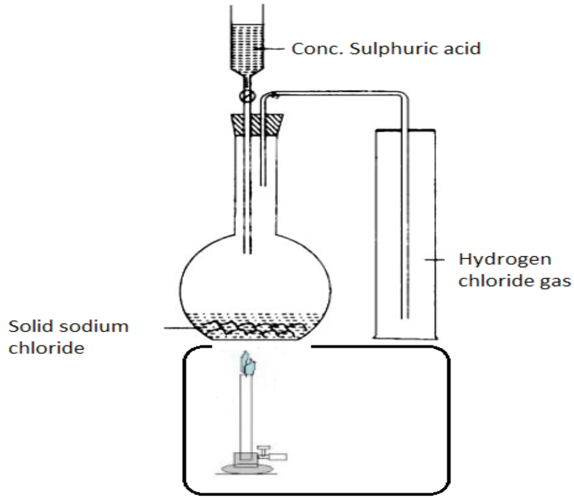
Apparatus

- Bunsen burner
- Wash bottle
- Funnel
- Round bottomed flask Nitric Acid
- Beaker Silver nitrate solution
- Glass rod Litmus paper
- Delivery tube

Chemicals

- Sodium chloride crystals
- Concentrated sulphuric acid
- Concentrated ammonia solution

Experiment setup:



Procedure:

1. Put about 5g of sodium chloride in round bottomed flask,
2. Set up the apparatus as shown above,
3. Using a dropping funnel, add slowly concentrated sulphuric acid to the sodium chloride,
4. Heat the mixture,

   What do you observe?

   Answer: *a colorless gas is released.*
5. Collect about 5 gas jars of the gas produced,

   i. **1st gas jar:** Remove the glass cover slip and try to pour out the gas into air.

   What happens?

   Answer: *on exposure to the air, the gas forms white fumes.*

   ii. **2nd gas jar:** Add a little water to the gas and shake. Test the resulting solution with litmus paper.

   What do you observe?

   Answer: *the solution turns red, blue litmus paper.*

iii. **3rd gas jar:** Add a little methylbenzene (toluene) to the gas and shake. Test the resulting solution with litmus paper or universal indicator.

Note what happens.

Answer: *no effect on the indicators. The gas dissolves in toluene, the gas does not ionize in toluene.*

iv. **4th gas jar:** Insert a glass rod that has been dipped in concentrated ammonia.

What do you observe?

Answer: *white fumes are observed*

v. **5th gas jar:** Insert a glass rod that has been dipped in silver nitrate solution.

What do you observe?

Answer: *a white precipitate is formed*

$HCl_{(g)} + NH_{3(g)} \rightarrow NH_4Cl_{(g)}$

White nitrate solution, formation of a white precipitate

$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

Interpretation of results and conclusion

Guiding questions

1. What happens when sodium chloride reacts with concentrated sulphuric acid? Write equation of the corresponding reaction
2. Describe the test of hydrogen chloride and write equation of the reaction when you test it.
3. How is hydrogen chloride tested and why?
4. What happens when hydrogen gas gets into contact with a silver nitrate? Write equation of the reaction that occurs

When sodium chloride reacts with concentrated sulphuric acid, hydrogen chloride gas and sodium hydrogen sulfate are formed. The equation of the reaction is:

$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$

Hydrogen chloride gas is collected by downward delivery as it is denser than air.

In the presence of water, it forms and acidic solution that turns red blue litmus paper.

$HCl_{(g)} \xrightarrow{Water} HCl_{(aq)}$ and $HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

The HCl is acidic.

In the presence of ammonia gas, dense white fumes are observed. These white fumes are ammonium chloride.

$HCl_{(g)} + NH_{3(g))} \rightarrow NH_4Cl_{(s)}$

When hydrogen gas gets into contact with a silver nitrate, a white precipitate of silver chloride is formed.

$AgNO_{3(aq)} + HCl_{(aq))} \rightarrow \underset{\text{White ppt}}{AgCl_{(s)}} + HNO_3(aq)$

Evaluation activity

1. What is the role of sulphuric acid contained in a wash bottle?
2. What is the difference between hydrogen chloride and hydrochloric acid?

Experiment 4.3: Reaction of diluted hydrochloric acid and metals.

Rationale: Generally, when dilute hydrochloric acid with metals, metal chlorides are formed. Many metal chlorides have many important uses: Sodium chloride (NaCl), also known as salt, is an essential compound our body uses to: absorb and transport nutrients.

KCl as a chemical feedstock, it is used for the manufacture of potassium hydroxide and potassium metal. It is also used in medicine, lethal injections, scientific applications, food processing, soaps, …

Calcium chloride has several similar uses as sodium chloride, and it is used as a food additive, food preservative and as brine in refrigeration plants. It is also used as a swimming pool chemical, in water treatment plants, and for desiccating purposes.

Objective: Learners will be able to carry out and explain the reaction of diluted hydrochloric acid with magnesium and zinc.

Required materials

Apparatus

- Test tubes
- Wooden splint
- Test tube holders

Chemicals

- Diluted hydrochloric acid
- Magnesium ribbon

Procedure

1. Put some pieces of magnesium ribbon in a test tube.
2. Add hydrochloric acid and test any gas evolved using a burning splint.

   Note down your observations.

   Answer: *a gas evolved burns with a pop sound.*

   Name this gas.

   Answer: *this gas is hydrogen*
3. Repeat step 1 and 2 using zinc and copper metals in place of magnesium.

   Note down your observations.

   Answer: *with zinc metal, a gas evolves and burns with a pop sound while with copper, no gas is evolved.*

Interpretation of results and Conclusion

Guiding questions

1. What happens when hydrochloric acid reacts with magnesium and zinc metals? Write equations of the corresponding reactions.
2. Does hydrochloric acid react with copper?

When hydrochloric acid reacts with magnesium metal, magnesium chloride and hydrogen gas are formed. The equation of reaction is:

$$Mg_{(s)} + 2HCl_{(g)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

Likewise, the reaction of hydrochloric acid with zinc yields zinc chloride and hydrogen gas.

$$Zn_{(s)} + 2HCl_{(g)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

The hydrochloric acid does not react with copper metal.

Experiment 4.4: Tests of chloride, bromide, and iodide ions in solution

Objective: Learners will be able to identify the halide ions by using dilute nitric acid, silver nitrate, and ammonia solution.

Required materials

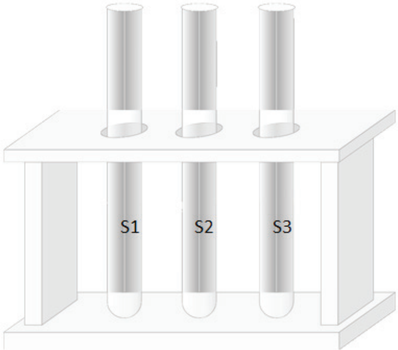
Apparatus

- Beakers
- Test tube holders
- Test tube rack
- Test tube
- Droppers
- Labels

Chemicals

- Sodium chloride solution
- Potassium iodide solution
- Potassium bromide
- Dilute nitric acid
- Silver nitrate solution
- Ammonia solution

Experiment set up



Procedure:

1. Prepare 3 test tubes and label them S1, S2 and S3,
2. In test tube S1 put 3 mL of sodium chloride; in test tube S2 put 3mL of potassium bromide and in test tube S3 put 3mL of potassium iodide,
3. To each test tube, add 1mL of dilute nitric acid followed by a few drops of silver nitrate solution. The nitric acid is added to remove other ions that might also give a confusing precipitate,

   Record your observations

   Observations:

Halide ions	Reactant	Observation
$Cl^-(aq)$	$AgNO_3(aq)$	White precipitate
$Br^-(aq)$	$AgNO_3(aq)$	Very pale-yellow precipitate
$I^-(aq)$	$AgNO_3(aq)$	Yellow precipitate

4. Divide the content of each test tube into two parts,
5. Then to one part add 5 drops of dilute ammonia solution and to the second part add 2 drops of concentrated ammonia solution.

Record your observations.

Observations:

Halide ions	Reactant	Observation
Cl^-	$NH_3(aq)$	The precipitate dissolves in as well dilute and concentrated ammonia solution to give colorless solution .
Br^-	$NH_3(aq)$	The precipitate is insoluble in dilute ammonia solution but dissolves in concentrated ammonia solution to give colorless solution.
I^-	$NH_3(aq)$	The precipitate is insoluble in both dilute and concentrated ammonia solution.

Interpretation of results and conclusion

Guiding question

1. Write ionic equation of reactions that occur when silver nitrate reacts with sodium chloride, potassium bromide and potassium iodide.

The ionic equation of reactions that occur when silver nitrate reacts with sodium chloride, potassium bromide and potassium iodide ions are:

$Ag^+(aq) + Cl^- \rightarrow AgCl(s)$ (white precipitate)

$Ag^+(aq) + Br^- \rightarrow AgBr(s)$ (pale-yellow precipitate)

$Ag^+(aq) + I^- \rightarrow AgI(s)$ (yellow precipitate)

The ammonia solution is used for confirming the presence of halide ions in solution.

UNIT 5 Rate of Reactions

Experiment 5.1: Effect of temperature on the speed of reaction

Rationale: The rate of reaction plays an important role in industrial processes in order to increase the productivity or reduce the cost of the product formed. The temperature is one of the factor which affect the rate of reaction. It is very crucial to investigate the effect of temperature on the rate of reaction and this will permit to increase the productivity in industry. The impact of temperature on the rate or reaction is an important factor in industrial processes. The temperature also is linked directly to the amount of energy required to the process, we know that the energy is expensive. We need to monitor the temperature which is appropriate to a specific experiment or process in order to increase productivity in industry.

Objective: Learners will be able to explain the effect of temperature on the speed of the react

Required materials

Apparatus

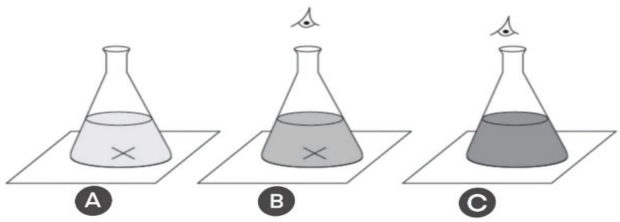
- 250 cm^3 conical flask
- Thermometer
- Stop clock
- Stand and clamp
- 50cm^3 measuring cylinder
- Gauze
- Graph paper
- Piece of white paper
- Test tubes
- 10 cm^3 measuring cylinder

Chemicals

- Sodium thiosulphate $Na_2S_2O_3$ (0.25M)
- Hydrochloric acid, HCl (2M)
- Distilled water

- Beaker (1L)
- $50cm^3$ conical flask
- Tripod

Experiment setup



Procedure

1. Using a $50cm^3$ measuring cylinder, put into a conical flask $10cm^3$ of the 0.25M sodium thiosulphate solution and add $40cm^3$ of distilled water.
2. Using a $10cm^3$ measuring cylinder, measure $5cm^3$ of the 2M hydrochloric acid and pour it into a test tube.
3. Prepare a water bath by half-filling a 1L beaker with water and place the beaker over a tripod stand and wire gauze.
4. Clamp the conical flask in the water bath and place the test tube in the water bath.
5. Place the thermometer in the conical flask and very gently heat the water bath until the contents of the flask reach the temperature of 30^oC.
6. Remove the conical flask and the test tube from the water bath.
7. Place the conical flask on the paper marked with a cross and immediately, add the acid from the test tube as you start the stopwatch. Gently stir the mixture with the thermometer.

   What do you observe?

   Answer: *A yellow solid forms in the solution.*
8. Observe the cross from above through the solution and stop the stopwatch as soon as the cross disappears from your view.

9. Record the time taken for the cross to be hidden and record the final temperature of the mixture.
10. Repeat the experiment for each of the temperatures of 25°C, 35°C, 40°C, 45°C, 50°C and record all of your results.

Data recording

Experiment	1	2	3	4	5	6
Temperature (°C)	25	30	35	40	45	50
Time taken for the cross to disappear (s)						

Interpretation of results and conclusion

Guiding questions

1. **What happened when** sodium thiosulphate $Na_2S_2O_3$ reacts with an acid?
2. What is the yellow precipitate formed?
3. Write equation of the reaction that occurs.
4. What happens when the temperature of the reactants is increased?

When sodium thiosulphate $Na_2S_2O_3$ reacts with an acid, a yellow precipitate of sulphur is formed. To follow this reaction, we can measure how long it takes for a certain amount of sulphur to form. This is done by observing how long it takes the black cross on white paper to be hidden.

The reaction that happens can be represented as follow.

$$Na_2S_2O_3(aq) + 2HCl\ (aq) \rightarrow 2NaCl\ (aq) + SO_2(aq) + H_2O\ (l) + S(s)$$

When the temperature of the reactants is increased, sulphur is produced more quickly as it takes less time for the cross to disappear. **This means that in the reaction the speed increases as the temperature increases.**

Note on source of error: A wrong recording of the time taken for the cross to disappear can lead to false results.

Evaluation activity

1. What are everyday examples of reactions whose rate is affected by temperature?
2. What is meant by saying that two quantities are directly proportional?
3. Suggest other factors that may affect the rate of the reaction.

Experiment 5.2: Effect of catalyst on the speed of reaction

Rationale: Catalyst is a substance that increases or decreases the rate of a chemical reaction by lowering the activation energy without itself being consumed by the reaction. Catalysts reduce the amount of energy required to break and form bonds during a chemical reaction. When the reaction is complete, catalysts remain chemically unchanged and they can be reused several times. The choice and being familiar with the use of catalysts in performing experiment or other everyday activities is the key in increasing productivity. Increasing of yield productivity in industries such as breweries, breads making among other involves the monitoring of the time of production and energy required in term of money spent due to long procedures. The choice of catalyst which is cheap and efficient is very needed in our daily life.

Objective: Learners will be able to explain the effect of the catalyst on the rate of the reaction

Required materials

Apparatus

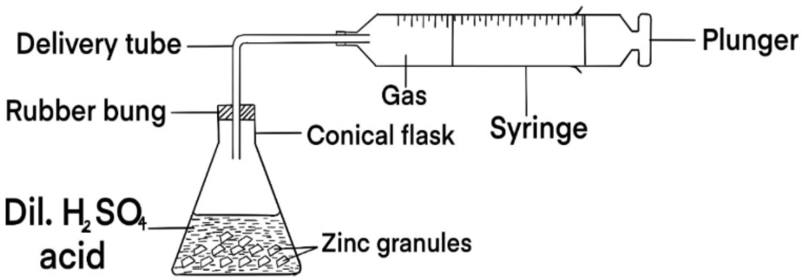
- 3 conical flasks
- Measuring cylinder (10 cm^3)
- Electronic balance
- Dropping pipette

Chemicals

- Zinc granules
- Dilute sulphuric acid (0.5M)
- Copper turning or powder
- Copper (II) sulphate solution

– Syringe
– Delivery tube
– A cork with one hole

Experiment setup: Reaction of zinc and hydrochloric acid



Procedure

1. Put approximately 5 grams of zinc granules into each of three conical flasks. Label them A, B and C.
2. Add 5 $cm^3$of dilute sulphuric acid to conical flask A. Measure the amount of gas produced as shown in the setup above.
3. In conical flask B, add a few copper turnings. Make sure they are in contact with the zinc granules. Add 5 $cm^3$of dilute sulphuric acid and using a cork with one-hole fitted with a delivery tube, stopper quickly the conical flask.
4. Measure the amount of gas produced as shown in the set up above.

   What do you observe?

   Answer: *more gas is produced.*

   Why are copper turnings added to zinc?

   Answer: *copper turnings are added to speed up the reaction.*
5. Add 5 $cm^3$of dilute sulphuric acid to conical flask C. Add about 1 $cm^3$of copper (II) sulphate solution using a dropping pipette. Quickly insert the cork with one-hole fitted with a delivery tube. Measure the amount of gas produced as in procedure 2 above.

What happens when copper sulphate is added to zinc?

Answer: *more gas is produced.*

6. Record all your results.

Data recording

Experiment	A	B	C
Amount of gas produced (in cm^3)			

Interpretation of results and conclusion

Guiding equations

1. Write equation of the reaction between zinc and dilute sulphuric acid.
2. What is the role of copper and copper sulphate in the reaction?

When pure zinc reacts with dilute sulphuric acid, the reaction is slow. Bubbles of hydrogen form on the surface of the zinc. The reaction that occurs is:

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

In conical flask B, copper turnings are the catalyst for the reaction and the reaction is faster than in conical flask A, but is not as fast as that in conical flask C.

In conical flask C, zinc displaces copper from the copper (II) sulphate solution and the surface of the zinc goes black. The displaced copper metal then acts as a catalyst for the reaction.

The reaction that happens is represented below.

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Evaluation activity

1. Explain how you would test the gas evolved in the reaction between zinc and dilute sulphuric acid.
2. Suggest other reactions that are speeded up by the presence of catalysts.
3. Do all reactions that require catalysts to proceed quickly? Illustrate your answer using at least three examples.

Experiment 5.3: Effect of concentration on the speed of reaction

Rationale: One of the factors that affect the rates of chemical reactions is concentration of the reactants, this factor is therefore used to increase the productivity in many chemical industries.

Objective: Learners will be able to explain the effect of the concentration on the speed of the reaction.

Required materials

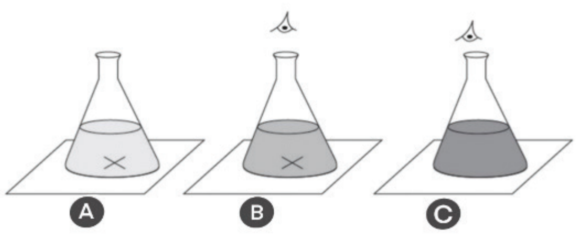
Apparatus

- Conical flask (250 cm^3)
- Measuring cylinder (100 cm^3)
- Stopwatch
- Piece of paper marked X

Chemicals

- 2M hydrochloric acid
- Sodium thiosulphate solution of different concentrations (0.06M, 0.12M, 0.18M, 0.24M and 0.30M)

Experiment setup



Procedure

1. Put 50cm^3 of 0.06M sodium thiosulphate solution into a conical flask.
2. Measure 5cm^3 of 2M hydrochloric acid into the measuring cylinder.
3. Add the acid to the conical flask and immediately start the stopwatch. Swirl the flask to mix the solutions and place it on a piece of paper marked with a cross.

4. Look down at the cross from above. When the cross disappears, stop the watch, and note the time taken. Record your results in the table shown below.
5. Repeat the procedure using a 0.12M, 0.18M, 0.24M and 0.30M

Data recording

Experiment	1	2	3	4	5
Concentration of Sodium thiosulphate in mol/dm^3					
Time taken for the cross to disappear					
(s^{-1})					

Interpretation of results and conclusion

1. What happens when sodium thiosulphate and dilute hydrochloric acid are mixed?
2. What is the effect of the concentration of n the rate of the reaction?

When sodium thiosulphate and dilute hydrochloric acid are mixed, a yellow precipitate of sulphur is formed. The higher the concentration of sodium thiosulphate solution, the less the time taken for the cross to disappear. **Increasing the concentration of one or more reactants will often increase the rate of reaction.**

Note on source of errors: A wrong recording of the time taken for the cross to disappear can lead to false results.

Evaluation exercise

1. What is the relationship between concentration and rate of reaction?
2. Does the concentration of a catalyst affect reaction rate?

Experiment 5.4: Effect of particle size on the rate of reaction

Rationale: One of the factors that affect the rates of chemical reactions is the particle of size of the reactants. this factor is therefore used to increase the productivity in chemical industries.

Objective: Learners will be able to explain the effect of particle size on the rate of the reaction.

Required materials

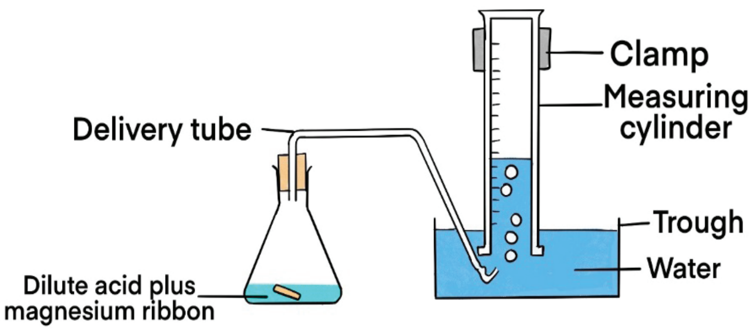
Apparatus

- Measuring cylinder ($100cm^3$)
- Stand and clamp
- 2 conical flasks ($100cm^3$)
- Weighing balance
- Glass trough
- Stopwatch
- Sand paper

Chemicals

- Magnesium ribbon
- Hydrochloric acid (1M)
- Distilled water

Experiment setup



Procedure

1. Clean the magnesium ribbon using a sandpaper to remove any oxides that may be coating its surface. This will reduce reaction errors related to impurities.
2. Cut three equal sizes (10cm each) of magnesium from the freshly cleaned magnesium ribbon, weigh each of them using a digital weighing balance and record their weights. Are they the same?
3. Wrap the magnesium pieces immediately in an aluminium foil to prevent them from being re-oxidised.
4. Fold the first magnesium ribbon (10 cm) and keep it for the next procedure.
5. Take the second magnesium ribbon cut it into smaller pieces of 1cm.
6. Take the third magnesium ribbon cut it into smaller pieces of 0.1cm.
7. Measure 20cm^3 of 1M hydrochloric acid using a clean dry measuring cylinder and pour into each of three clean 100cm^3 conical flasks and label them A, B and C.
8. Successively put the folded magnesium ribbon 10 cm in A, the fine pieces of magnesium 1cm in B and 0.1cm magnesium ribbon in conical flask C. Immediately start the stopwatch.
9. Monitor the reaction progress closely and stop your stopwatch when the magnesium completely dissolves in the acid or reaction comes to stop. In each case, record the reaction duration in seconds in the table below.

Data recording

	A	B	C
Time taken for the complete reaction(s)			

Interpretation of results and conclusion

Guiding questions

1. Write equation of the reaction that occurs when magnesium reacts with hydrochloric acid
2. Compare the rate of reaction when sulphuric acid reacts with fine pieces or large pieces of magnesium.

The reaction involving magnesium in very fine pieces (0.1 cm) takes fewer time to came to an end than the reactions involving magnesium in larger pieces (1 and 10 cm). Also, effervescence was more rapid with very fine pieces. The reaction that occurs is represented by the following chemical equation.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Decreasing the size of the particles, which make up a given weight, will increase the number of particles represented by the same weight. **Smaller particle size results in an increase in the speed of reaction because the surface area of the interacting reactants has been increased.**

Evaluation

1. How does the particle size affect the speed of the reaction?
2. What would happen if the piece of magnesium ribbon was ground into powder?
3. Suggest at least two examples in the daily life illustrating the effect of particle size on the reaction speed.

Experiment 5.5: Effect of light on the rate of reaction

Objective: Learners will be able to explain the effect of light on the rate of the reaction.

Required materials

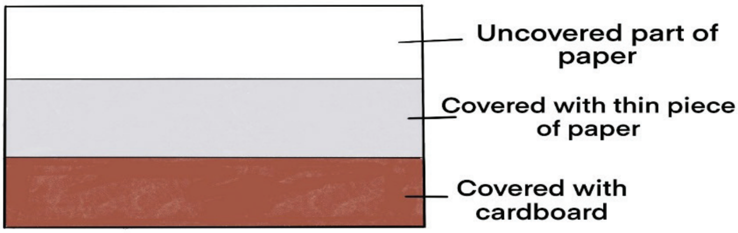
Apparatus

- Pair of scissors
- Cardboard
- Dark room
- Art brush
- A4 paper

Chemicals

- Silver bromide solution

Experiment setup



Caution: Silver bromide is harmful as it may irritate eyes and the skin.

Procedure

1. Wear protective gloves and safety goggles.
2. Cut a piece of paper into a square shape using a pair of scissors then cut a piece of cardboard to the same size as paper.
3. In a dark room smear the paper with silver bromide solution.
4. Cover the lower part of the paper with cardboard as shown in the setup above.
5. Cover the upper part with a thin piece of paper.
6. Expose the improvised film to light; leave it for about 5 minutes.
7. Remove all the coverings then make your observations.

Interpretation of results and conclusion

Silver bromide decomposes to silver metal which will be seen as black spots. On the cardboard area no black spots seen because light cannot pass through it.

Silver bromide ($AgBr$), silver chloride ($AgCl$) and silver iodide (AgI) are all sensitive to light and they are used in the production of various types of photographic films.

In this reaction silver ions are reduced to silver metal and the bromide ion is oxidised to bromine molecule.

$Ag^{+}(aq) + e- \rightarrow Ag(s);$

$2Br(aq)^{-} \rightarrow Br_2(l) + 2e^{-}$

Evaluation activity

1. Using proper chemical equations, give two examples of reaction that are sensitive to light.
2. How does light affect the vision?

UNIT 6 Chemical Properties of Acids snd Bases

Experiment 6.1: Reaction of water and sodium metal

Objective: Learners will be able to predict and describe a reaction between water and metal.

Required materials

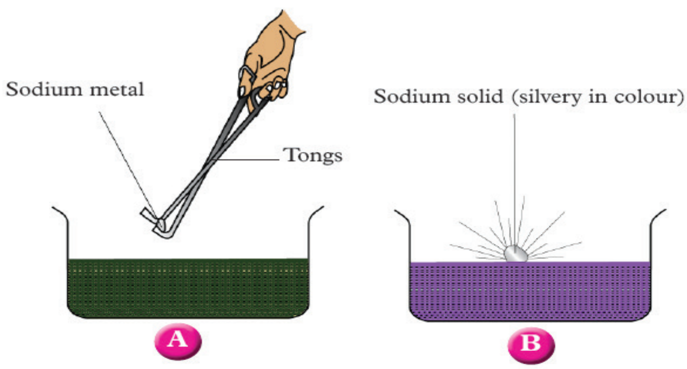
Apparatus

- Through
- Splint
- Bunsen burner
- Test tube Universal indicator
- Knife
- Pair of tongs
- Wire gauze

Chemicals

- Sodium metal
- Water
- Aluminium foil

Experiment setup



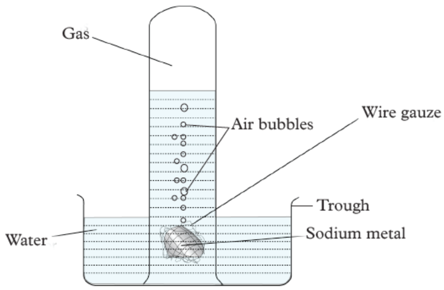
Procedure:

1. Place water in a trough up to half its volume,
2. Add three drops of phenolphthalein to the water,
3. Cut a small piece of sodium (size of rice grain) and drop it into the water in the trough,

   Note the observations.

   Answer: *Sodium moves quickly around on the surface of the water. A colourless gas is given off and the solution turns pink.*

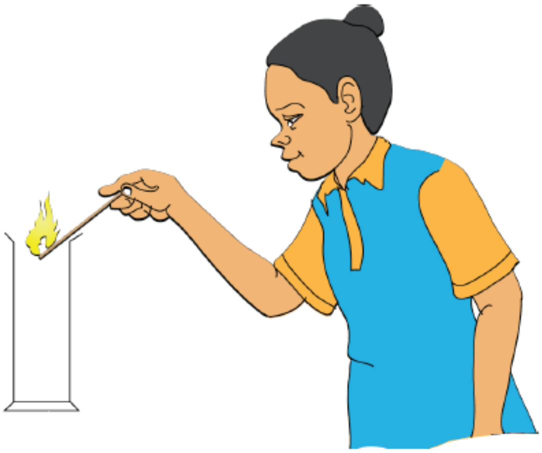
4. Wrap another piece of sodium in wire gauze. Place it in a trough of water and insert on it a test tube full of water as shown. Collect the gas in the test tube.



5. Test the collected gas in the test tube with a burning splint.

   Note your observations.

   Answer: *the gas burns with a 'pop' sound.*



What is this gas?

Answer: *this gas is hydrogen*

Interpretation of results and conclusion:

Guiding questions

1. Write a chemical equation of the reaction between water and sodium metal.
2. Why sodium floats on water?
3. Why the hissing sound is produced?

Sodium floats on water because it is less dense than water. It reacts with water to produce a gas that burns with a 'pop' sound. The gas is hydrogen. It moves swiftly on the water surface since the hydrogen gas evolved is trying to escape from all directions. A hissing sound is produced due to the heat generated during the reaction.

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

The phenolphthalein turns pink indicating the solution formed is alkaline (sodium hydroxide).

Evaluation

1. Why is sodium wrapped in wire gauze?
2. Write a chemical equation of the reaction between water and calcium metal.
3. What the role of the indicator in this reaction?
4. What would you observe if universal indicator were used in the place of phenolphthalein?

Experiment 6.2: Reaction of diluted hydrochloric acid with metals, metal carbonates, metal hydrogen carbonates and metal oxides

Objective: Learners will be able to illustrate the reaction of diluted hydrochloric acid with metals, metal carbonates, metal hydrogen carbonates and metal oxides

Required materials

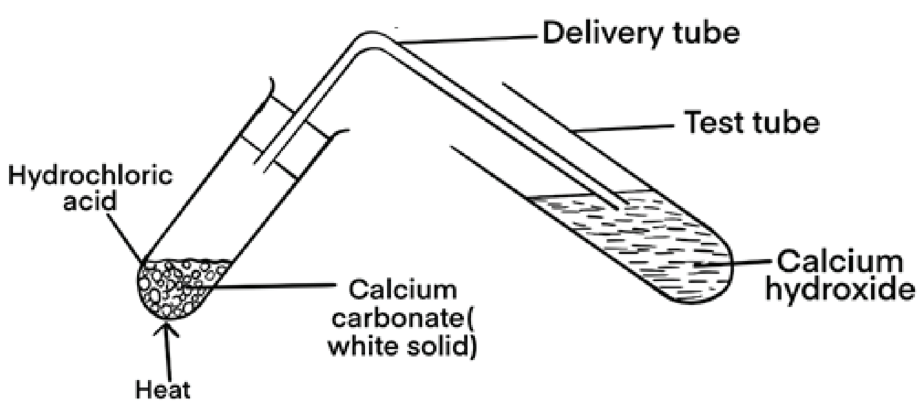
Apparatus

- Test tubes
- Splint
- Bunsen burner
- Beakers
- Test tube rack
- Test tube holders
- Delivery tube
- Stand and clamp

Chemicals

- Dilute hydrochloric acid
- Magnesium metal
- Calcium carbonate
- Sodium hydrogen carbonate
- Magnesium oxide
- Sodium hydroxide
- Phenolphthalein indicator

Experiment setup



Procedure:

a) Reaction of dilute hydrochloric acid with magnesium

1. Place 5 cm^3 of dilute hydrochloric acid in a test tube. To the acid, add a piece of magnesium ribbon,

   Record your observations.

   Answer: *a colourless gas is evolved*

2. Introduce a burning splint in the test tube,

   What do you observe?

   Answer: *the colourless gas burns with a 'pop' sound`*

   What is this gas?

   Answer: *this gas is hydrogen,*

b) Reaction of dilute hydrochloric acid with calcium carbonate and sodium hydrogen carbonate

1. Place 5 cm^3 of dilute hydrochloric acid into a test tube. Add half endful spatula of calcium carbonate,
2. Insert a burning splint into the test tube where the reaction is occurring,
3. Repeat procedure 1 using sodium hydrogen carbonate in place of calcium carbonate,

   What do you observe?

   Answer: *effervescence occurs, and a colourless gas is liberated. The gas puts off a burning splint.*

   What is this gas?

   Answer: *this gas is carbon dioxide*

c) Reaction of dilute hydrochloric acid with magnesium oxide.

1. To 5 cm^3 of dilute hydrochloric acid in a test tube, add magnesium oxide and heat,

   What do you observe?

   Answer: *the magnesium oxide dissolves.*

Interpretation of results and conclusion

Guiding question

1. Write chemical equations of the reactions that occur when hydrochloric acid reacts with metals, carbonates, hydrogen carbonates and metal oxides.

   a) When magnesium reacts with dilute hydrochloric acid, magnesium chloride and hydrogen gas are produced. The equation of the reaction is:

   $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(g) + H_2(g)$

A similar reaction occurs when dilute sulphuric acid is reacted with magnesium metal.

   b) When calcium carbonate or sodium hydrogen carbonate reacts with dilute hydrochloric acid effervescence occurs and a colourless gas is liberated. The gas puts off a burning splint. The gas is carbon dioxide. The equations of the reactions are:

   $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

   $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$

   c) When a mixture of magnesium oxide and hydrochloric acid is heated, the magnesium oxide dissolves reacting with the acid to form magnesium chloride and water.

   $MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$

Evaluation activity

1. Write a chemical equation of the reaction of dilute sulphuric acid with calcium.
2. Why was there heating in the reaction of dilute hydrochloric acid with magnesium oxide?
3. What would happen if methyl orange was used in place of phenolphthalein?

UNIT 7 Concentration of Solutions

Rationale: Concentrations are used for chemical reactions and informs you how much you have, in how much volume. A chemical reaction that requires specific amounts of reagents may just not work without the proper concentrations. In some cases not knowing the concentration of some reagents can incur risk with strong acids and bases. Another great example is pharmaceuticals. Concentration is very important, and is related to the prescribed dosage of medicine. The wrong concentration can be a serious issue.

Concentration is directly tied to determining information like how much heavy metal is in water, or how much pesticide ran off into a nearby water source, or how strong an instance of acid rain is. The list goes on. Without concentration usage, these types of analysis would be much more difficult or even impossible.

Experiment 7.1: Preparation of solutions

a) Preparation of 1M solution of NaCl

Objective: Learners will be able to prepare a 1M sodium chloride solution

Required materials

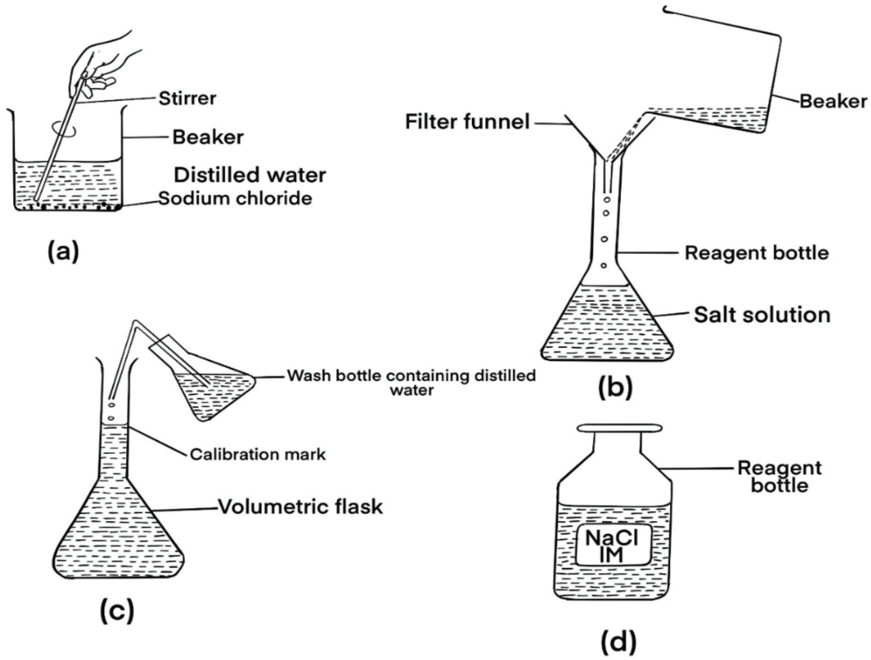
Apparatus

- Weighing balance
- Washing bottle
- Glass rod stirrer
- Beaker
- Funnel

Chemicals

- Distilled water
- Sodium chloride

Experiment setup



Procedure:

1. Using the relative atomic masses, find the molar mass of sodium chloride.

   RAM: Na =23, Cl =35.5

   What is the mass required to prepare a sodium chloride solution of 1mole/L?

   Answer: *(23+ 35.5) g =58.5g*

2. Weigh the mass of sodium chloride needed 58.5g, as it calculated in step 1, to prepare 1M NaCl solution.
3. Place 400mL of distilled water in a beaker.
4. Add 58.5g sodium chloride into the water, stir continuously until it is completely dissolved.
5. Using a filter funnel transfer the salt solution into a 1L volumetric flask.
6. Rinse the beaker with distilled water and transfer carefully washings into the flask.

7. Add more distilled water and shake flask well. Add more water until the solution level is just below the calibration mark.
8. Using a washing bottle add distilled water drop by drop until the bottom of the meniscus is at the same level with calibration mark.
9. Use the stopper to cover the flask and invert it several times to make sure the solution is well mixed.
10. Transfer the solution into a reagent bottle and label it 1M NaCl.

b) Preparation of 500ml of NaOH 0.1M

Procedure:

1. Find molar mass of NaOH: Mm NaOH= 23+16+1= 40g/mol
2. NaOH 0.1M means 0.1 moles of NaOH per litre of solution
3. In 500ml there are (0.1x500)/1000= 0,05mol.
4. Mass of NaOH to be dissolved in water to make 500ml of solution= 40g x 0,05= 2g.
5. Add 2g sodium hydroxide into the water, stir continuously until it is completely dissolved.
6. Using a filter funnel transfer NaOH solution into a 500 mL volumetric flask.
7. Rinse the beaker with distilled water and transfer carefully washings into the flask.
8. Add more distilled water and shake flask well. Add more water until the solution level is just below the calibration mark.
9. Using a washing bottle add distilled water drop by drop until the bottom of the meniscus is at the same level with calibration mark.
10. Use the cork to cover the flask and invert it several times to make sure the solution is well mixed.
11. Transfer the solution into a reagent bottle and label it 1M NaOH.

Interpretation of results and conclusion

The total volume of the solute plus the solvent must be equal to 1000mL, therefore a 1L volumetric flask is used.

When one mole of a solute is dissolved in solvent (water) and the volume of the solution is made up to 1L, the solution is said to be a molar solution.

A concentration of one mole per litre is often written as 1mol/L or 1 mol. dm^{-3} or 1M. Concentration expressed this way is sometimes referred to as the molarity of a solution, abbreviated as M. Therefore, a molar solution contains 1 mole of solute per $1dm^3$ solution.

Evacuation activity

1. How many grams of calcium chloride are needed to make 250 mL of a 0.50 M solution?
2. Explain how to make at least 0.5 litre of a 1.25 molar ammonium hydroxide solution.
3. How many grams of ammonia are present in 2.0 L of a 0.1M solution?

Experiment 7.2: Titration of hydrochloric acid with sodium hydroxide

Objective: Learners will be able to determine the concentration of a base or an acid using titration.

Required materials

Apparatus

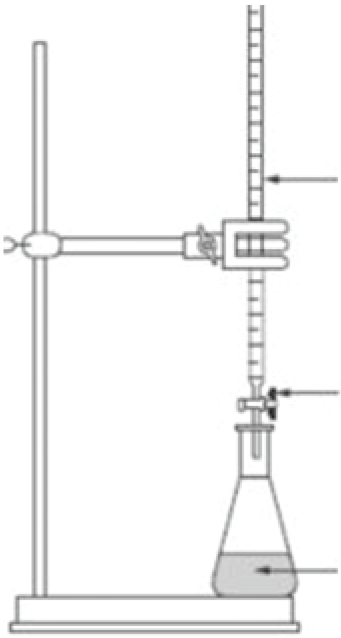
- Retort stand
- Clamp
- Burette
- Pipette
- Conical flask
- Beaker

Chemicals

- Standard solution of NaOH
- HCl with unknown concentration
- Phenolphthalein or methyl orange
- Distilled water

- Wash bottle
- White tile or paper
- Funnel
- Dropper
- Measuring cylinder

Experiment setup



Procedure

1. Pipette 25.0mL of sodium hydroxide.
2. Using a funnel, transfer the sodium hydroxide into a conical flask.
3. Add 3 drops of phenolphthalein indicator as shown on the diagram above.

   What happens?

   Answer: *The solution becomes pink*

4. Using a filter funnel, fill the burette with a solution of hydrochloric acid until it is up exactly to zero mark.

5. Remove the filter funnel.
6. Place a white tile or paper under the conical flask to clearly see the colour changes.
7. Add dropwise the hydrochloric acid to the sodium hydroxide solution and swirl after each addition of acid for thorough mixing while at the same time keeping an eye on the solution in the flask to notice any colour change.
8. Add the acid until the alkali is completely neutralized. This is shown by colour change of content of the conical flask.
9. When you swirl the flask and the new colour developing persists for a while, the end point is near. Add one drop of the acid at a time until you get a permanent colour. This is the **endpoint** of the titration.
10. Read and record the volume of the acid to the nearest 0.1mL. Record your results in a table like the one shown below.
11. Repeat the titration two more times using a different clean conical flask. Add the same number of drops of the indicator as used in the first titration. However, if you use the same flask, you must wash it thoroughly with distilled water and then rinse with a little of the solution you intend to suck into the pipette.

    Note: The first volume obtained should guide you to get the second volume fairly quickly as follows:

    "If your first titre volume was 25.0 cm^3, you may add 21 cm^3 quickly while swirling the flask. Then add the acid dropwise while swirling the conical flask until you get the same permanent colour change as in previous titration."
12. Repeat the titration using the second volume as a guide to the volume required. You can add the acid until you are about 1mL from the end point, then you add the acid drop by drop.
13. Get the average of any values for the volume of the acid added, which differ by not more than ± 0.1mL. In other words, the three titrations should be fairly consistent and accurate.

Data recording

Burette reading	**1(Trial)**	**2**	**3**	**4**
Final burette reading (mL)				
Initial burette reading (mL)				
Volume of acid used (mL)				

Interpretation of results and conclusion

1. Calculate the average volume of hydrochloric acid added to reach the end point using two concordant values.
2. Use this value to calculate the molar concentration of sodium hydroxide.

Evaluation activity

Perform the titration of potassium hydroxide with sulphuric acid using methyl orange indicator

UNIT 8 Electrolysis and its Applications

Rationale: Electrolysis is one of processes with many important uses. For example, it's used in the extraction of metals from their ores. Another example it is used for refining. Certain metals, such as copper and zinc, can be used for the manufacture of chlorine, and it can be used for electroplating materials that we use on a day to day basis.

Experiment 8.1: Electrolysis of water

Objective: Learners will be able to perform and explain the electrolysis of acidified water.

Required materials

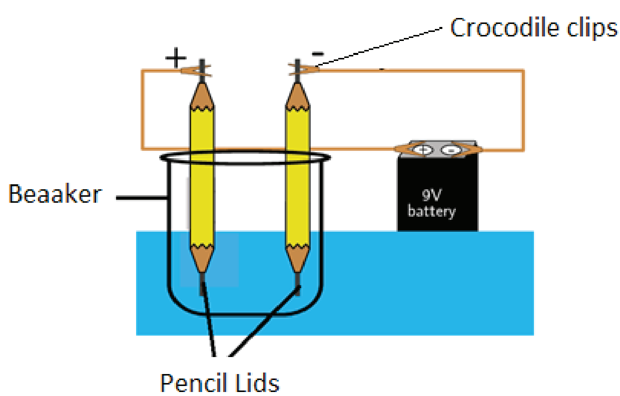
Apparatus

- Beaker
- Two alligator clip leads or insulated wire
- 6-volt or 9-volt battery
- Two sharpened pencils or graphite rods

Chemicals

- Acidified water (aqueous
- sulphuric acid)

Experiment setup



Procedure

1. Fill the beaker with acidified water.
2. Fill two test tubes with acidified water.
3. Hold two test tubes full of water above each electrode.
4. Set up the apparatus as shown above and connect the electrolyser to the source of electricity.

   What do you observe after a while?

   Answer: *bubbles of gas are evolved around each electrode.*

Interpretation of results and conclusion

Guiding questions

1. During electrolysis of water, what happens to hydrogen ions (H^+) and hydroxide ions (OH^-)?
2. Write equations that are showing reactions at the cathode and anode and equation of overall reactions.

During electrolysis of water, the hydrogen ions (H^+) are attracted to the cathode where they are reduced to hydrogen molecules and hydroxide ions (OH^-) to the anode where they are oxidised to oxygen molecules.

Reaction at cathode (negative electrode)

$$2H^+(aq) + 2e^- \rightarrow H_2\,(g)$$

Reaction at anode (positive electrode)

$$4OH^-\,(aq) \rightarrow O_2\,(g) + 2H_2O\,(l) + 4e^-$$

Overall reaction

$$2H_2O\,(l) \rightarrow 2H_2(g) + O_2(g)$$

The volume of hydrogen is double the volume of oxygen.

Note that in case of electrolysis of tape water, ions present would compete with the OH^- and H^+ from the decomposition of water and may lead to false results.

Evaluation activity

1. Why is a small amount of sulphuric acid added to water in this process?
2. If electrolysis of water produced 8 ml of hydrogen, what were the volume of oxygen produced?

Experiment 8.2: Electrolysis of molten lead bromide

Objective: Learners will be able to perform and explain the electrolysis of lead (II) bromide.

Required materials

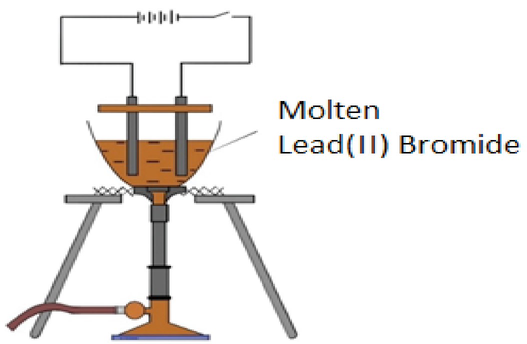
Apparatus

- Crucible
- Two alligators clip leads or insulated wire
- 6-volt or 9-volt battery
- Two carbon rods
- Bunsen burner
- Match box
- Triangle

Chemicals

- Lead (II) bromide

Experiment setup



Caution: Bromine vapours are toxic. Work in fume hood.

Procedure:

1. Wear safety glasses, gloves, and protective masks.
2. Weigh 10g of lead (II) bromide and put it in a crucible.

   What is the colour of lead (II) bromide?

   Answer: *the colour of lead (II) bromide white*

3. Set up the apparatus as shown above
4. Switch on the electric current.

   What do you observe?

   Answer: *electricity does not pass throuth the solid.*

5. Heat strongly the crucible until the lead (II) bromide melt.

   Why is lead (II) bromide melted?

   Answer: *to make free its ions.*

6. Keep heating the crucible for five minutes.

   What do you observe?

   Answer: *grey solid is deposited to the cathode and brown fumes appear at the anode.*

Interpretation of results and conclusion

Guiding questions

1. Why electrolysis is not possible with solid lead (II) bromide but possible with molten lead bromide?
2. Write equations of the reactions at electrodes

Electrolysis is not possible with solid lead (II) bromide. This is because the ions are held in a three-dimensional lattice, unable to move freely to the electrodes. Melting enables the ions to become mobile and to travel to the respective electrodes.

Thus, molten lead bromide, $PbBr_2$(l) is an electrolyte. During electrolysis, Pb^{2+} **ions are attracted to the cathode where they gain electrons and get reduced to lead atoms, (Pb)**. Br^- ions are attracted to the anode where they lose electrons and get oxidised to bromine molecules (Br_2).

Reaction at electrodes

At anode: $2Br^- \rightarrow Br_2(g) + 2e^-$

At cathode: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

Grey globules of lead are observed at negative electrode whereas brown fumes of bromine are observed at the positive electrode.

Evaluation activity

1. Is electrolysis of lead (II) bromide a redox reaction? Explain.
2. Why is potassium not extracted by electrolysis of its aqueous salt solution?
3. Explain why a graphite anode is preferred to other inert electrodes during electrolysis of fused lead bromide.

Experiment 8.3: Electrolysis of concentrated sodium chloride solution

Objective: Learners will be able to perform the electrolysis of concentrated sodium chloride solution (2M) and identify the products.

Required materials

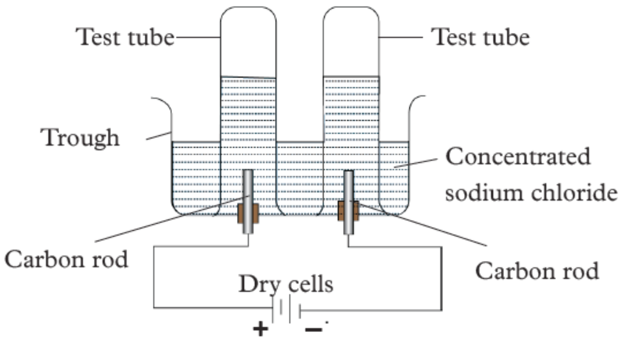
Apparatus

- Electrolyser/beaker 2M
- 6-volt or 9-volt battery or dry cells
- Match box
- Wooden splint
- Test tubes
- Carbon rods

Chemicals

- Sodium chloride

Experiment setup:



Procedure

1. Set up the apparatus as shown above.
2. Half-fill the electrolyser with 2M sodium chloride solution.
3. Switch on the current.

   What is the colour of the gas collected at the anode and the cathode?

   Answer: *Gas at the anode is green-yellow and the gas at the cathode is colourless.*

4. Put moist litmus papers in the test tube of the anode.

   What do you observe?

   Answer: *the gas bleaches moist litmus papers and therefore the gas is chlorine.*

   At which electrode is a gas with a smell produced?

   Answer: *Anode*

5. Test the gas produced at each electrode with a burning splint.

   At which electrode does the gas burn with a "pop" sound?

   Answer: *Cathode*

   Which gas that is produced at the cathode?

   Answer: *Hydrogen*

Interpretation of results and conclusion

Guiding questions

1. What do you observe at the electrodes when the circuit is completed?
2. Which gases are collected at the anode and the cathode?
3. Write equations of the reactions that occur on the electrodes.

When the circuit is completed, bubbles are observed at the electrodes. A green-yellow poisonous gas which has a choking irritating smell is collected at the anode. This gas bleaches moist litmus papers and therefore the gas is chlorine.

At the cathode, a colourless gas which burns with a 'pop' sound is collected. This gas is therefore hydrogen.

Reactions at electrodes

At the anode: At this electrode, $Cl^-(aq)$ and $OH^-(aq)$ are present but although OH^- ion is lower in the electrochemical series than Cl^-, the concentration of Cl^- ions at anode is greater than the OH^- ions.

$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$

At the cathode: $Na^+(aq)$ and $H^+(aq)$ are present but H^+ ions are discharged in preference to Na^+ ions because H^+ ion is lower in the electrochemical series than Na^+ ion.

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

H^+ and Cl^- ions are discharged leaving Na^+ and OH^- ions. Finally, sodium chloride solution is converted to sodium hydroxide solution.

Experiment 8.4: Electrolysis of dilute sodium chloride solution (1M)

Objective: Learners will be able to perform and explain the electrolysis of dilute sodium chloride solution.

Required materials

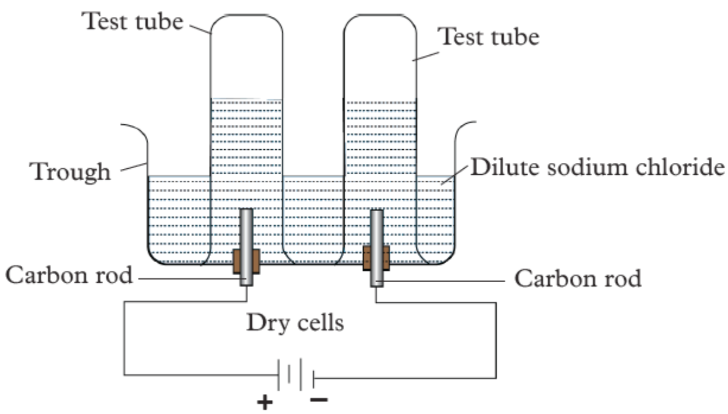
Apparatus

- Beaker/ electrolyser
- 6-volt or 9-volt battery or dry cells
- Match box
- Wooden splint
- Test tubes
- Test tubes holder
- Stirrer
- Carbon rods

Chemicals

- Dilute sodium chloride (1M)
- Distilled water

Experiment setup



Procedure

1. Set up the apparatus as shown in the figure above
2. Half-fill the electrode with a dilute solution of sodium chloride.
3. Switch on the current.

   What do you observe on the electrodes?

   Answer: *Colourless gases are evolved. The level of sodium chloride in the test tube decreases.*

4. Test the gases collected with a glowing and burning splints.

What do you observe?

Answer: *At the anode a colourless gas re-lights a glowing splint. This gas should be oxygen. At the cathode, a colourless gas burns with a 'pop' sound. This gas should be hydrogen.*

Interpretation of results and conclusion

Guiding questions

1. Which gases are collected at electrodes?
2. Write equations of reactions that occur on the electrodes and overall reaction.

A colourless gas, which re-lights a glowing splint, is collected at the anode. The gas is therefore oxygen. A colourless gas which burns with a 'pop' sound is collected at the cathode. This gas is therefore hydrogen.

Reactions at electrodes

At the anode: OH^- and Cl^- ions are attracted to the anode, OH^- ions give up to the anode to form water and oxygen gas while Cl^- ions remains in solution.

$$4OH^-(aq) \rightarrow 2H_2O\ (l) + O_2\ (g) + 4e^-$$

At the cathode: H^+ and Na^+ ions are attracted to the cathode, H^+ ions gain electrons from the anode to form hydrogen gas. Na^+ ions remain in solution.

$$2H^+\ (aq) + 2e^- \rightarrow H_2(g)$$

Overall reaction:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

The electrolysis of dilute sodium chloride is equivalent to the electrolysis of water. Two volumes of hydrogen are liberated for one volume of oxygen.

Since water is being removed during electrolysis, the concentration of sodium chloride solution increases gradually.

Experiment 8.5: Electrolysis of copper (II) sulphate using graphite electrodes

Objective: Learners will be able to carry out and explain electrolysis of copper (II) sulphate using inert electrodes.

Required materials

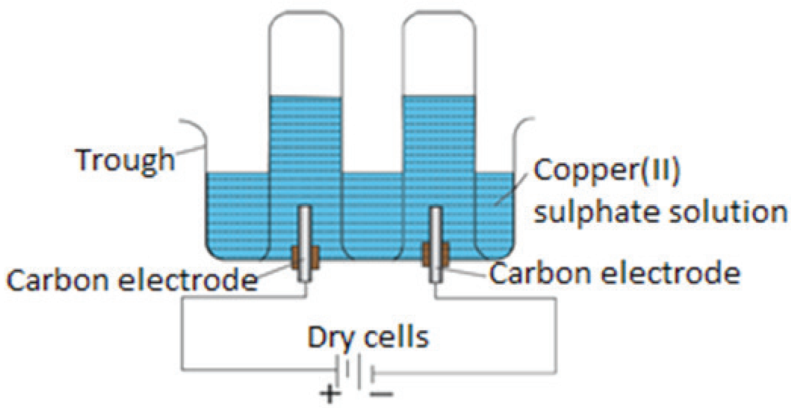
Apparatus

- Electrolyser (with carbon electrodes)
- Battery/dry cells
- Connecting wires fitted with crocodile clips
- Match box
- Wooden splint

Chemicals

- Copper (II) sulphate solution

Experiment setup



Procedure

1. Half-fill an electrolyser with copper (II) sulphate solution.
2. Set up the apparatus as shown in the figure above.

   What is the purpose of the bulb?

   Answer: *to verify if copper (II) sulphate solution conduct electricity and therefore can be electrolysed.*

3. Switch on the current and after a short while, observe what happens on the carbon electrodes.

Answer: *At the anode a colourless gas is evolved; at the cathode a reddish sold is deposited.*

Interpretation of results and conclusion

Guiding questions

1. Write equations of the reactions at the electrodes.
2. Why the intensity of the blue colour of the electrolyte decreases?

At cathode, Cu^{2+} and H^{+} ions are presents.

Cu^{2+} ions are selectively discharged to form copper metal.

$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$

At anode, OH and ions are presents.

OH- ions are selectively discharged to form oxygen gas.

$4OH^{-}(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^{-}$

The intensity of the blue colour of the electrolyte decreases because the concentration of copper (II) sulphate solution decreases as more copper is deposited on the cathode and if we continue with electrolysis the solution finally turns colourless and acidic. A colourless gas, which re-lights a glowing splint, is collected at the anode. The gas is therefore oxygen.

Evaluation activity

1. At the end of electrolysis, the solution finally turns colourless and acidic. Explain this.
2. If we continue passing electricity in the colourless solution what would be the product at: (i) cathode? (ii) anode?

Experiment 8.6: Electrolysis of copper (II) sulphate using copper electrodes

Objective: Learners will be able to perform and explain the electrolysis of copper (II) sulphate using copper electrodes

Required materials

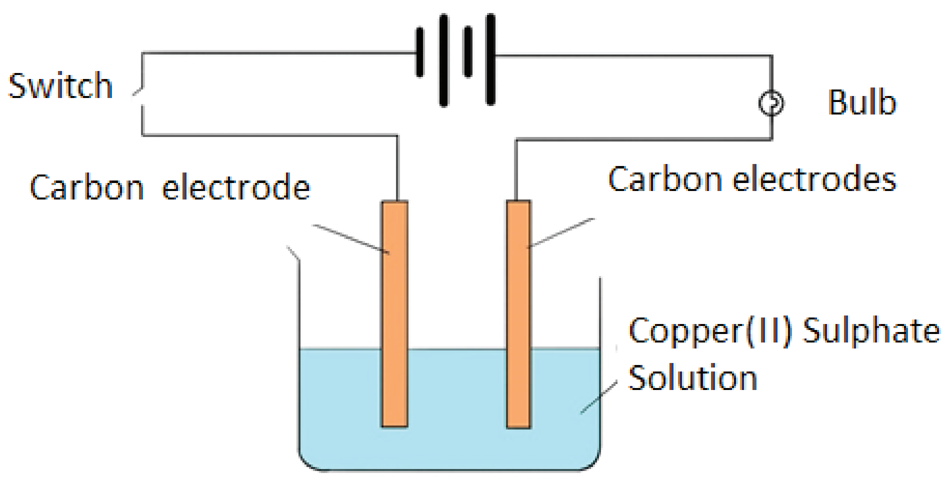
Apparatus

- Beaker/trough
- Battery/dry cells Distilled water
- Copper rods (2)
- Connecting wires fitted with crocodile clips
- Bulb (6V)
- Electronic balance

Chemicals

- Copper (II) sulphate solution
- Propanone

Experiment setup



Procedure

1. Put copper (II) sulphate solution in a beaker in figure above.
2. Clean the copper electrodes until they shine.

   Why must electrodes be cleaned before electrolysis?

   Answer: *To remove any oxide that might have formed on the surface.*

3. Weigh the electrodes and record the mass in your notebook as in table below.
4. Set up the apparatus as shown in figure above and switch on electric current.
5. After 10 minutes, note the colour of the electrolyte and rinse the electrodes with distilled water.
6. Dip the electrodes in propanone and then let them dry off.

   Why must the electrodes be washed with distilled water and propanone after electrolysis?

   Answer: *The electrodes are washed with water and propanone to ensure they are completely dry.*

7. When dry, re-weigh the electrodes.
8. Record the mass obtained after the experiment.

Data recording

	Before electrolysis		**After electrolysis**	
	Anode	Cathode	Anode	Cathode
Colour of the copper (II) solution				
Mass of electrode				
Change in mass of electrode				

Interpretation of results and conclusion

Guiding questions

1. Write equations of the reactions that occur at electrodes.
2. Explain why the blue colour of the solution remains unchanged during the electrolysis.

Reactions at electrodes

At the cathode: Cu^{2+} and H^+ ions are attracted to the cathode. Cu^{2+} ions are selectively discharged, giving copper metal (pink solid).

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

H^+ ions remain in solution

At the anode: S and OH^- ions are attracted to the anode. Copper anode dissolves, giving Cu^{2+} ions.

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

S and OH^- ions remain in solution.

The overall reaction is the transfer of copper from the copper anode to the copper cathode.

The masses of both electrodes change during electrolysis. The anode loses mass while the cathode gains mass. That is, during electrolysis, the copper anode slowly becomes smaller, and the copper cathode slowly becomes bigger as pure copper is deposited on it.

The concentration of aqueous copper (II) sulphate remains unchanged. Thus, the blue colour of the solution remains unchanged during the electrolysis as Cu^{2+} ions that are discharged at the cathode are constantly replaced by more Cu^{2+} ions formed from the copper atoms at anode.

Evaluation activity

1. What is the purpose of electrolysing copper sulphate solution using copper electrodes?
2. What would happen if a silver coin were kept in a solution of copper sulphate for few days?

Experiment 8.7: Electroplating of silver coin using copper (II) sulphate

Rationale: There are metals that are easily attacked by corrosion and others have an appearance that needs to be improved, thus electroplating is a durable solution to deal with the above-mentioned issue in non- electroplated materials.

Objective: Learners will be able to perform and explain electroplating of a silver objects using copper (II) sulphate solution.

Required materials

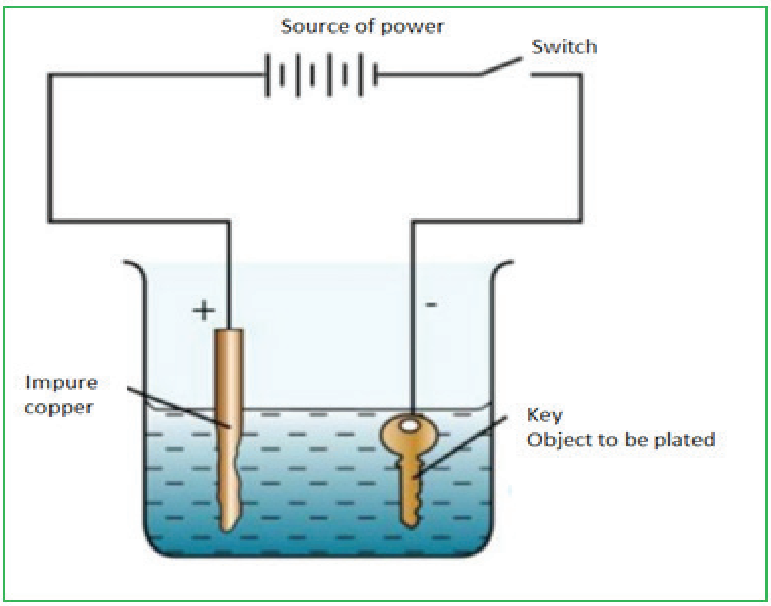
Apparatus

- Two crocodile clips
- 1.5volt DC battery with battery holder
- Copper electrode
- A key
- Beaker
- Safety equipment

Chemicals

- Copper (II) sulphate solution

Experiment setup



Procedure

1. Prepare the key for copper-plating by cleaning it with toothpaste or soap and water. Dry it off on a paper towel.
2. Stir copper sulphate into some hot water in a beaker until no more dissolves. The solution should be dark blue. Let it cool.
3. Use one crocodile clip to attach the copper electrode to the positive terminal of the battery (this is now the anode) and the other to attach the key to the negative terminal (now called the cathode)
4. Suspend the key in the solution of copper sulphate
5. Place the copper electrode into the solution, making sure it does not touch the key and the solution level is below the crocodile clip.
6. Leave the circuit running for 20-30 minutes.

Data recording

	Anode	Cathode
Ions present		
Reaction taking place		
Colour of the coin before electroplating		
Colour of the coin after electroplating		

Interpretation of results and conclusion

Guiding questions

1. Write equations of the reactions that occur at electrodes.
2. Explain what happens at the cathode and the anode.

The coating of a metal object with another metal object is called **electroplating**. The object to be plated is made the cathode in the cell so that metal ions move to it when inserted in electrolyte.

Reaction at cathode (negative electrode)

$Cu^{2+}(aq) + 2e^- \rightarrow Cu\ (s)$ Copper metal deposit at cathode.

Reaction at anode (positive electrode)

Copper atoms dissolve from anode, copper ions pass in the solution

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

Copper metal of anode will dissolve and added in electrolyte. Before the electrolysis the silver coin has a grey-white colour. After the electrolysis the key has a brown colour as it is covered by a layer of copper metal. Direct current should be used.

Evaluation activity

1. Are electroplated objects alloys? Explain.
2. What will happen if a silver rod is left in a copper nitrate solution for few days?
3. Mention at least three applications of the electroplating.

UNIT 9 Structure and Properties of Alkenes and Alcohols

Experiment 9.1: Preparation and test for ethene

Rationale: Preparation of ethene is one of the most important reactions as it has many uses in every day life. In medicine it is used as an anaesthetic. Ethene is as used as oxy-fuel gas in metal cutting. It is used also as refrigerant, used in the manufacture of polythene (polymer). Ethene is also used to form other useful compounds like ethylene glycol (an antifreeze agent), ethanol etc

Objective: Learners will be able prepare and test for the presence of ethene (C=C).

Required materials

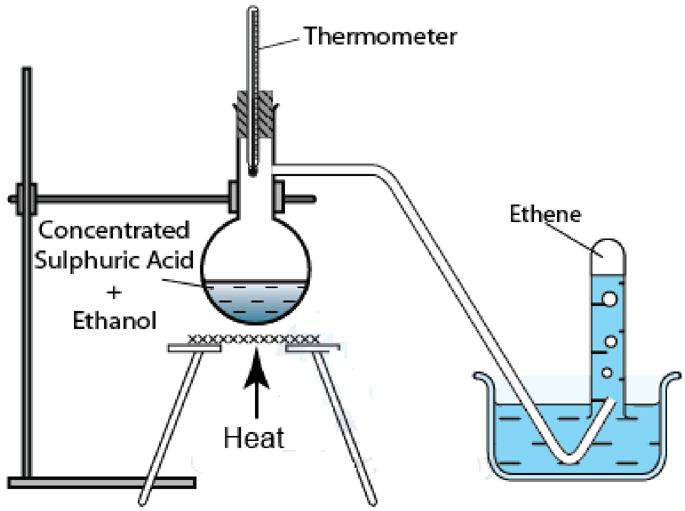
Apparatus

- Test tubes
- Thermometer
- Gas jar
- Tripod stand
- Retort stand
- Cork with one hole

Chemicals

- Concentrated sulphuric acid
- Ethanol

Experiment set up



Procedure

1. Arrange the apparatus as shown in figure above and the mixture of concentrated sulphuric acid and ethanol to temperatures between 170^0-180^0C.

   Note your observations

   Answer: *A colourless gas is collected into a gas jar over water.*

2. Bubble ethene through acidified manganate (VII) solution.

   What do you observe?

   Answer: *Acidified manganate (VII) solution is decolorized.*

3. Bubble ethene through bromine water.

   What do you observe?

   Answer: *Bromine water is decolorized*

4. Bubble ethane through bromine water and acidified manganate (VII) solution

   What do you observe?

   Answer: *bromine water and acidified manganate (VII) solution are not decolorized*

5. What can you conclude from the above observations?

   Record your observations and conclusions.

Interpretation and Conclusion

Guiding questions

1. To test for ethene or alkenes in general, bromine water is used. Describe this identification test.
2. Write down equation of the reaction that occur when ethene/ethylene reacts with bromine.

To test for ethene or alkenes in general bromine water is used. An alkene turns brown bromine water colorless as it reacts with the double bond. Bromine water remains brown in the presence of an alkane as there is no double bond

Equation of the reaction

$$C_2H_{4(g)} + Br_{2(aq)} \rightarrow C_2H_4Br_{2(aq)}$$

Colorless Brown Colorless

When ethene (an alkene) reacts with potassium manganate (VII), it is also oxidized. Thus, bromine water and acidified potassium manganate (VII) are used to confirm the presence of an alkene.

Experiment 9.2: Fermentation of alcohol

Rationale: In everyday life alcohols are used as solvents in marker pens, medicines, and cosmetics (such as deodorants and perfumes).

Beer and wine are produced by fermenting glucose with yeast. Yeast contains enzymes that catalyze the breakdown of glucose to ethanol and carbon dioxide. In this experiment, a glucose solution is left to ferment. Alcohol has a crucial need in everyday life. Its demand in local beverages, medicine and in hair salon is high.

Objective: Learners will be able to carry out fermentation efficiently using banana juice and yeast.

Required materials

Apparatus

- Conical flask
- Delivery tubes
- Trough
- Cork with one hole
- Cotton wool

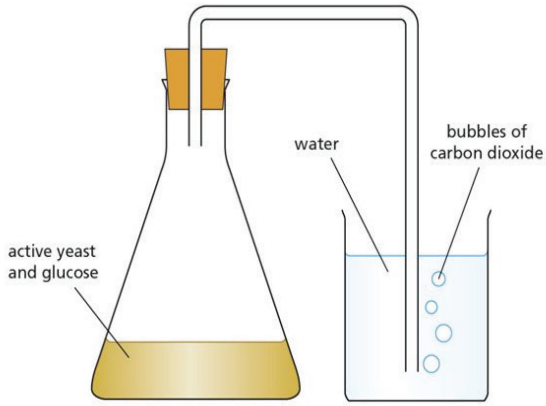
Chemicals

- Sugar
- Water
- Yeast
- Limewater

Procedure

1. Pour 50 cm^3 of warm water in a conical flask then add 5 g of glucose. Swirl the flask to dissolve the glucose.
2. Add 1 g of yeast to the solution and loosely plug the top of the flask with cotton wool.
3. Wait the fermentation to take place.
4. Remove the cotton wool and pour the invisible gas into the boiling tube containing limewater. Take care not to pour in any liquid as well.
5. Gently swirl the limewater in the boiling tube and note what happens.
6. Replace the cotton wool in the top of the flask.
7. Remove the cotton wool and note the smell of the solution.

Experiment set up



$$C_6H_{12}O_6 \xrightarrow{\text{Zymase}} 2\ CH_3CH_2OH \quad + 2CO_2$$

Interpretation of results

Alcoholic fermentation is a biotechnological process accomplished by yeast, some kinds of bacteria, or a few other microorganisms to convert sugars into ethyl alcohol/ethanol and carbon dioxide. Yeast contains enzymes that catalyse the breakdown of glucose to ethanol and carbon dioxide. In this experiment, a glucose solution is left to ferment.

$$C_6H_{12}O_6 \xrightarrow{\text{Zymase}} 2\ CH_3CH_2OH \quad + 2CO_2$$

During alcoholic fermentation as alcohol is formed, carbon dioxide released increased proportionally. The end of fermentation is marked by the stop release of carbon dioxide. Alcohol is widely used as solvent, fuel, beverage and it is used as component of hand sanitizer, disinfectants, paints …

UNIT 10 Carboxylic Acids

Experiment 10.1: Reaction of carboxylic acids with sodium carbonate, sodium hydroxide, alcohol, and sodium metal

Rationale: Carboxylic acids occur in many common household items: Vinegar contains acetic acid, aspirin is acetylsalicylic acid, vitamin C is ascorbic acid, lemons contain citric acid, and spinach contains oxalic acid. Therefore, having deep knowledge in properties of carboxylic acids in much needed.

Objective: Learners will be able to show how carboxylic acids react with sodium carbonate, sodium hydroxide, alcohols, and sodium metal.

Required materials

Apparatus

- Test tubes
- Droppers
- Beaker
- Spatula
- Test tube holder
- Source of heat

Chemicals

- Calcium carbonate
- Sodium metal
- Ethanol
- Ethanoic acid
- Concentrated sulphuric acid
- Sodium hydroxide solution
- Phenolphthalein indicator

Procedure

1. Place four test tubes in a test tube rack and label them.
2. Add about 2mL of ethanoic acid in each test tube.
3. Add a small piece of sodium in the first test tube and record your observation. Test for gas produced if any.

   What do you observe?

   Answer: *A colourless gas which burns with a 'pop' sound is observed.*

4. Add a spatula end full of calcium carbonate in the second test tube and record your observation. Test for gas produced if any.

   What do you observe?

   Answer: *A colorless gas which turns milky lime water is observed.*

5. Add an equal amount of ethanol followed by two drops of concentrated sulfuric acid in the third test tube.

6. Warm gently the mixture in hot water and smell the product formed and record your observation.

   What do you observe?

   Answer: *The formed product has a fruity smell.*

7. Add two drops of phenolphthalein indicator in the fourth test tube.

8. Add 2mL of sodium hydroxide solution to the solution drop by drop until there is a color change.

   What do you observe?

   Answer: *The color changes from colorless to pink.*

Interpretation of results and conclusion

Guiding question

1. Write equations of the reactions when carboxylic acids react with sodium metal, sodium carbonate, ethanol and sodium hydroxide.

Carboxylic acid reacts with carbonates, metals, bases, and alcohol to give different products. Ethanoic acid reacts with metals to produce hydrogen gas which burns with pop sound.

$$CH_3COOH\ (aq) + Na\ (s) \rightarrow CH_3COONa\ (aq) + H_2\ (g)$$

Ethanoic acid reacts with sodium carbonate to produce carbon dioxide which turns milky lime water.

$$2CH_3COOH\ (aq) + Na_2CO_3\ (s) \rightarrow 2CH_3COONa\ (aq) + CO_2\ (g) + H_2O\ (l)$$

Ethanoic acid reacts with ethanol in presence of concentrated sulphuric acid to produce ethyl ethanoate which is an ester with fruity smell.

$$CH_3COOH\ (aq) + CH_3CH_2OH\ (l) \xrightarrow{\text{Conc. } H_2SO_4} CH_3COOCH_2CH_3\ (aq) + H_2O\ (l)$$

Ethanoic acid react also with sodium hydroxide in neutralization reaction to produce sodium acetate salt.

CH_3COOH (aq) + NaOH (aq) → CH_3COONa (aq) + H_2O (l)

Evaluation activity

Complete the following equations:

a) CH_3CH_2COOH (aq) + Na (s) →

b) CH_3CH_2COOH (aq) + CH_3OH (l) →

c) CH_3CH_2COOH (aq) + KOH (l) →

d) CH_3CH_2COOH (aq) + KOH (l) →

e) CH_3CH_2COOH (aq) + $MgCO_3$ (s) →

Experiment 10.2: Esterification

Rationale of the experiment

Derivative products of esters are very common in our everyday life. Perfumes and cosmetics, soaps and detergents are ester derivatives and they are so important. Esters are highly needed in different domains. Understanding the esterification process will help manufacturer to get employees who can make enough quality ester substances which are crucial raw materials in different manufacturing processes.

Objective: Learners will be able to carry out and explain esterification experiment.

Required materials

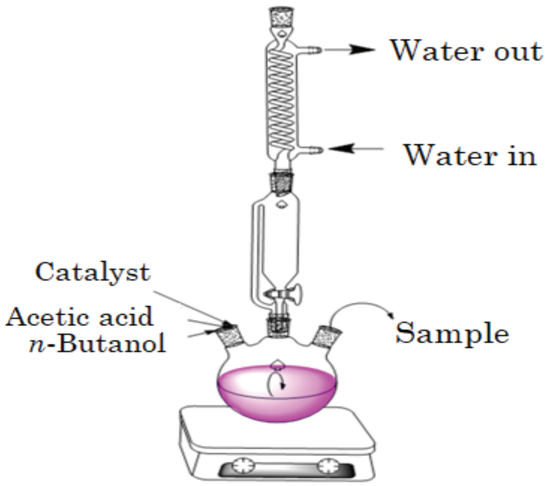
Apparatus

- Liebig condenser
- Thermometer
- Hot plate
- Funnel

Chemicals

- Cold water
- Alcohol/ butanol
- Diluted sulphuric acid
- Acetic acid

Experiment set up



Esterification setup

Procedure

1. Set apparatus as it is shown in the figure above
2. Pour 5cm^{3}of n-butanol and 6 cm^3 acetic acid
3. Heat the mixture on a hot plate to a temperature of about 60^{0}C
4. Leave mixture on the hot plate for 15 minutes
5. Cool the mixture by immersing it in cold water
6. Detect odours with caution because vapors produced may be harmful.
7. Did you feel any smell change?

   Answer: *a sweet smell is formed.*

Interpretation of results and conclusion

Guiding questions

1. Write equation of the reaction that occur in this experiment.

The reaction of combining an organic acid (RCOOH) with an alcohol (ROH) to form a compound called ester (RCOOR) and water. This reaction is called esterification. Ester is obtained by an esterification reaction of an alcohol and a carboxylic acid.

Equation of the reaction:

$$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2OH + CH_3CH_2COOH \rightleftharpoons CH_3\text{-}CH_2\text{-}COOCH_2CH_2CH_2CH_3 + H_2O$$

Evaluation activity

Complete the following equations of reactions:

a) CH3OH + CH3COOH

b) CH3CH2OH + CH3CH2CH2COOH

Experiment 10.3: Preparation of soap (Saponification) and detergents

Rationale of the experiment

Soaps and detergents are very common in our life. We are used to buy them from different shops however they can be made in our laboratories and even at home. Nowadays, the cost of soaps and detergents is increasing. After this lesson, students will be familiar with the process to make their own soaps at home.

Objective: learners will be able to make soap via saponification reaction

Required materials

Apparatus

- 250 mL beaker
- Glass stirring rod
- Glass pipets and pipet bulbs
- pH paper
- Test tubes
- Tube rack
- 10 mL graduate cylinder
- Magnetic stir bar
- Hot/stir plate

Chemicals

- Olive oil
- Vegetable oil or coconut oil
- Food colouring
- 6 M sodium hydroxide solution
- Assorted fragrances(optional)
- Stearic acid

Procedure:

1. Heat the olive oil or vegetable oil or coconut oil to 35°C in a 150mL beaker.
2. Pour 10 mL of the warm oil into a tall 250 mL beaker.
3. Choose your fragrance. You may choose one of the following: holiday candy, island coconut, lavender, cinnamon, vanilla.
4. Add 1-2 drops of desired fragrance, using the pipet provided at front bench; do not mix fragrances.
5. Add 3 mL of 20% sodium hydroxide solution to the beaker. This is approximately two full dropper squirts.
6. Use the glass stirring rod to mix. You must stir for 20-45 minutes; you may choose to take turns with your lab partner. The mixture will slowly become smoother and opaquer.
7. Add 2-3 drops of desired food coloring and continue to stir.
8. Add 1/8 teaspoon or the tip of a spatula of stearic acid to harden the liquid soap. Stir.
9. Pour into chosen mold shape. Label with your names.

10. After pouring into the mold (small beaker), the process will continue on its own. The soap will heat up and liquefy again, then cool off slowly, harden and dry. So, the soap must be left undisturbed for at least 12 hours. You will pick up your finished soap in lab next week.

Interpretation of results

Guiding questions

1. What is the name of the process for making soap?
2. What is the role of sodium hydroxide?
3. State the importance of fragrance
4. Give the importance of using food colouring in the experiment
5. In the experiment, stearic acid is added. State the role of stearic acid.

A solid soap with its characteristics including forming foam/lather with water is formed.

Soap is made by reacting fat and hydroxide in the process called **Saponification. Saponification** is a process that produces soap, form fats and sodium hydroxide. Saponification involves base hydrolysis of triglycerides, which are esters of fatty acids, to form the sodium salt of a carboxylate(Soap).

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