REB | RWANDA BASIC EDUCATION BOARD

CHEMISTRY EXPERIMENTS USER GUIDE

FOR

SENIOR FOUR

FOREWORD

Dear teacher,

Rwanda Basic Education Board (REB) is honoured to present Chemistry experiments user guide for Senior four. 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 use 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, AIMS – TTP in collaboration with Mastercard Foundation who provided technical and financial support and REB staff particularly those from the Mathematics and Science Subjects Unit in the Curriculum Teaching and Learning Resources Department. Special appreciation goes also to teachers and independent experts in education who supported the exercise throughout the process. Any comment or contribution would be welcome for the improvement of this booklet for next versions.

Dr. MBARUSHIMANA Nelson

Director General, REB

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

Special thanks are given to, AIMS staff, IEE staff, URCE-Lecturers, 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 Mathematics and Science subjects Unit in the CTLR Department who were involved in the whole process of the development work of this user guide.

Finally, my word of gratitude goes to AIMS – TTP in partnership with MasterCard Foundation for their support in terms of human and financial resources towards the development of this guide which will strengthen STEM teaching hence improving the quality of education in Rwandan schools.

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
KBC	:	Knowledge Based Curriculum
UR-CE	:	University of Rwanda- College of Education
CTLR	:	Curriculum, Teaching and Learning Resources
REB	:	Rwanda Basic Education Board
AIMS-TTP	:	African Institute for Mathematical Sciences-Teacher Training Program

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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 of chemicals and apparatuses 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:

- Analyses 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 the 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 rationle, 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:**

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 the 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,
- Analyse 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	It can lead to death
 Irritant	It 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	It can easily explode and releases small particles which can injure you
 Harmful	Poisonous when inhaled and ingested and can lead to death

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 reactions, 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 labelled 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 WORKSHEET

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 the teacher or be available 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/ observation	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

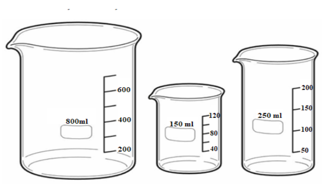
After conducting a laboratory experiment, students should write a report about their findings and the conclusion they reached. The report to be made depends on the level of students. The following is a structure of the report to be made by a group of secondary school learners (S1-S3).

1. Introduction (details related to the experiment: Students identification, date, year, topic area, unit title and lesson).
2. The title of the 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 for 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
2	Beaker		Used to hold and heat liquids. Multipurpose and essential in the laboratory.
3	Brushes		Used to easily clean the inside of a test tubes and other glassware.
4	Buchner funnel		Used with vacuum flask for performing vacuum filtration.
5	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.

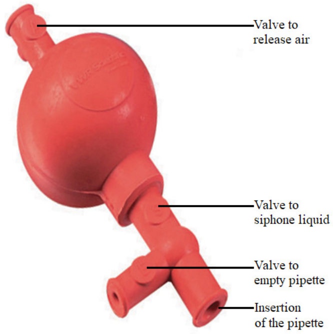
6	Burette		Operating the Burette Proper burette technique is an important laboratory skill that may take some practice to develop. Although it may seem initially awkward, a right handed person should operate the burette with the left hand, and a left handed person should operate the burette with the 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 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.

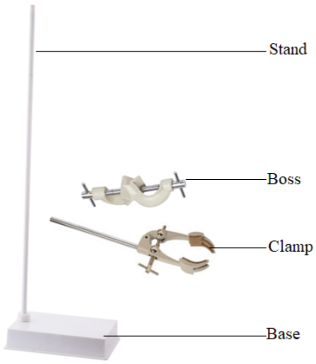
7	Burette clamp		Used to hold burettes on a ring stand.
8	Clay triangle		Used to hold crucibles when they are being heated. They usually sit on a ring stand.
9	Crucible with lid		Used to heat small quantities to very high temperatures.
10	Crucible tong		Used to hold crucibles and evaporating dishes when they are hot.
11	Disposable pipette		Used for moving small amounts of liquid from place to place. They are usually made of plastic and are disposable.

12	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.

13	Erlenmeyer flasks/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.
14	Evaporating dish		Used to recover dissolved solids by evaporation.
15	Forceps		Used for picking up and moving small objects.
16	Glass funnel & Polypropylene funnel		Used to pour liquids into any container so they will not be lost or spilled. They are also used with folded filter paper for filtration.
17	Glass stir rod		Used to stir liquids. They are usually made of glass.

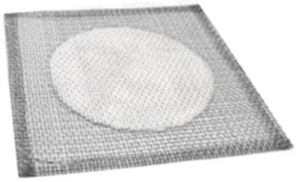
18	Graduated cylinder/ measuring cylinder		Used to measure the volumes of liquids.
19	Micropipette		Used for accurately measuring and delivering very small volumes of liquid-usually 1 ml or less. Steps to follow when using a micropipette Select the volume. Set the tip. Press and hold the plunger at the 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.
20	Mortar and pestle		Used to crush solids into powders for experiments, usually to better dissolve the solids.

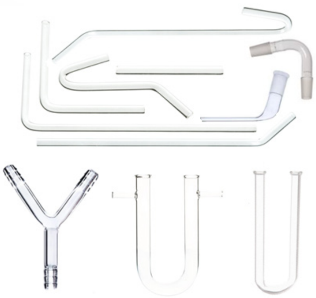
21	Pipette filler		How does a pipette filler work? Siphon liquid into the pipette to the desired level by squeezing valve "S" on the bottom of the pipette filler. This uses the vacuum created in the bulb to draw liquid into the pipette. Be careful not to draw liquid into the pipette filler. ... This allows you to release liquid at the desired rate and to the desired level.
22	Pipette with pump		Used for accurately measuring and delivering small volumes of liquid- usually 0.1-10 mL.
23	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.

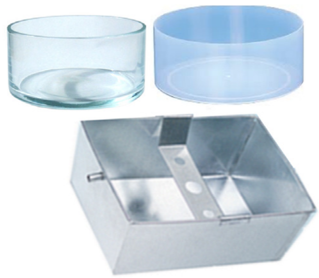
24	Retort stand and accessories		Used to hold items being heated. Clamps or rings can be used so that items may be placed above the lab table for heating by Bunsen burners or other items. Used also to hold burette
25	Rubber stopper		Stoppers come in many different sizes. The sizes are from 0 to 8. Stoppers can have holes for thermometers and for other probes that may be used.
26	Separator funnel		For separating layers of immiscible liquids or for dropping liquids.
27	Spatula		Used for moving small amounts of solid from place to place.

28	Test tube		Used for storing, mixing, and heating small amounts of chemicals.
29	Test tube holder		Used to hold test tubes while heating.
30	Test tube rack		Used to hold test tubes while reactions happen in them or while they are not needed.
31	Thermometer		Used to take temperature of solids, liquids, and gases.

32	Utility clamp		Used to attach test tubes and other glassware to retort stand.
33	Vacuum filter flask		Used with vacuum line and Buchner funnel for vacuum filtration.
34	Volumetric flask		Used to prepare solutions with accurate concentration.
35	Wash bottle		Used to wash; rinse containers

36	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.
37	Wire gauze		Used with a ring clamp to support glassware over a Bunsen burner.
38	Borosilicate glass tube		Used to connect to other items of glassware or equipment to deliver chemicals, solvents, liquids, gases and other products.
39	Deflagrating spoon or gas jar spoon		Generally used for the burning of materials inside gas jars or similar.

40	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.
41	Cardboard cover		
42	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.
43	Borosilicate delivery tube		The delivery tube is particularly useful for bubbling a gas from a gas cylinder or stoppered vessel through a liquid.

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

50	Charcoal rod		Charcoal rods are used as electrodes
51	Syringe		They are often used for measuring and transferring solvents and reagents where a high precision is not required.
52	Electrolyser		Electrolyser is used in the electrolysis process
53	Gas jar and cover		A container used for collecting gas from experiments.

B. List of chemicals

SN	Chemicals
1	Sodium chloride
2	Potassium chloride
3	Distilled water
4	Sugar
5	Metallic rod (iron rod)
6	Sulphur rod
7	Aluminium foil
8	Iron rod
9	Sulphur rod
10	Charcoal rod
11	Sodium metal
12	Potassium rod
13	Sodium oxide
14	Red litmus paper
15	Blue litmus paper
16	Sodium hydroxide
17	Potassium hydroxide
18	Phenolphthalein
19	Lithium carbonate
20	Sodium carbonate
21	Potassium carbonate
22	Lime water
23	Lithium nitrate
24	Sodium nitrate
25	Potassium nitrate
26	Sodium sulphate
27	Potassium sulphate
28	Lithium hydroxide
29	Rubidium hydroxide
30	Lithium carbonate

31	Lithium chloride
32	Potassium sulphate
33	Potassium chloride
34	Sodium sulphate
35	Sodium chloride
36	Magnesium ribbon
37	Magnesium rod
38	Calcium rod
39	Hydrochloric acid
40	Calcium oxide
41	Calcium carbonate
42	Magnesium nitrate
43	Magnesium hydroxide
44	Calcium hydroxide
45	Magnesium sulphate
46	Barium hydroxide
47	Barium sulphate
48	Aluminium foil
49	Hydrochloric acid
50	Sulfuric acid
51	Potassium permanganate
52	Aluminium powder
53	Aluminium oxide
54	Aluminium hydroxide
55	Aluminium salt solution
56	Ammonia solution
57	Charcoal
58	Carbon powder
59	Tin rod
60	Tin Oxide
61	Tin (II) chloride
62	Lead rod

63	Lead Oxide
64	Lead (II) chloride
65	Silicon oxide
66	Silicon (II)chloride
67	Mercury nitrate
68	Calcium hydroxide
69	Red phosphorus
70	Nitric acid
71	Copper rod
72	Copper turnings
73	Zinc powder
74	Sulphur powder
75	Bromine
76	phosphoric acid
77	Phosphorus oxides
78	Phosphorus chloride
79	Calcium phosphate
80	Potassium dichromate
81	Sodium sulphite
82	Ammonium molybdate
83	Potassium chloride
84	Potassium bromide
85	Potassium iodide
86	Phosphate salts
87	Hydrated copper (II) sulphate
88	Hydrated iron (II) sulphate
89	Iodine
90	Sodium thiosulphate
91	Sodium sulphite
92	Sodium sulphate
93	Barium nitrate
94	Sodium bromide

95	Sodium/Potassium iodine
96	Sodium chloride
97	Starch
98	Silver nitrate
99	Potassium chloride
100	Zinc rod
101	Copper Nitrate

PART 3: DETAILED EXPERIMENTS FOR SENIR FOUR

UNIT: 3

FORMATION OF IONIC AND METALLIC BONDS

EXPERIMENT 3.1: ELECTRICAL CONDUCTIVITY OF IONIC COMPOUNDS

Rationale

Electrical conductivity of solutions may help to measure the strength of electrolytes. The knowledge on the conductivity of some solutions help to prevent accidents which can happen when such solutions are in contact with electric wires and when one is dealing with plants containing sap in direct contact with electric wires. This experiment is performed to illustrate the conductivity of ionic compounds' solutions.

Objective

Learners will be able to carry out experiments and explain the electrical conductivity of aqueous solution of sodium chloride and ionic compounds in general.

Required materials

Apparatus

– An electric circuit with batteries
– Electric bulb
– Conducting wires
– Crocodile clips
– Graphite electrodes (2)
– Beakers

Chemicals

– Distilled water
– Sodium chloride
– Potassium chloride
– Sugar

Experimental set-up

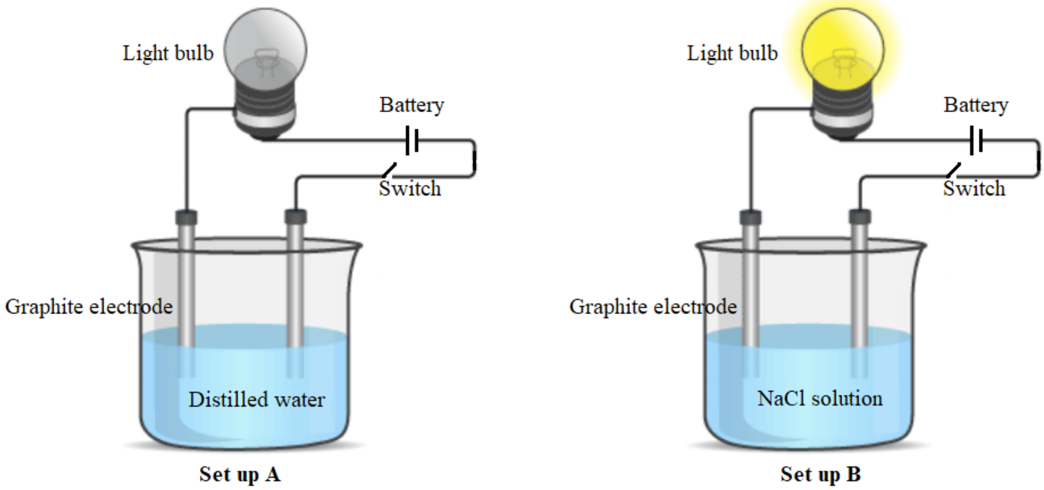


Figure 3.1: Electrical conductivity of distilled water and NaCl solution

Procedure

1. Set the apparatus as shown in Figure 3.1.
2. Switch on the circuit of both set up A and B. What do you observe?
3. Repeat the same experiment using potassium chloride and sugar solutions and note your observations.

Answer: *The bulb does not give light when connected in distilled water and sugar solution, but it lights up when connected in NaCl and KCl solutions.*

Interpretation of results and conclusion

Guiding question

Explain your observations.

Answer to guiding question

– Distilled water does not conduct electricity. Therefore, it is a non-electrolyte.

– Sodium chloride (NaCl) and potassium chloride (KCl) solutions conduct electricity because they ionise into Na^+ and Cl^- and K^+ and Cl^- respectively in water. Thus, they are electrolytes.

Most ionic compounds such as sodium chloride and potassium chloride are ionised in water. Their aqueous solutions conduct electricity. Therefore, they are called electrolytes. Aqueous solution of sugar does not conduct electricity. Thus, it is not electrolyte. It does not ionise in water. Substances like sugar which do not ionise in water and do not conduct electricity are called non-electrolytes.

Evaluation

1. Classify the following substances under electrolytes and non-electrolytes and justify your answer.
    a) Magnesium bromide
    b) Hexene
    c) Potassium chloride
    d) Sugar
    e) Ethanol
    f) Sulphuric acid
    g) Acetic acid

Answer: *Electrolytes: Magnesium bromide, Potassium chloride, Sulphuric acid and acetic acid are electrolytes as they are ionised into ions in water*

Non-electrolytes: Hexene, sugar and ethanol are non-electrolytes as they are not ionised in water

2. Carry out the experiment to investigate the conductivity of solution X containing potassium chloride and solution Y containing sugar. Interpret your observation.

EXPERIMENT 3.2: CONDUCTIVITY OF HEAT BY METALS

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	Chemicals
– Bunsen Burner	– Metallic rod (iron rod)
– Match box	– Sulphur rod

Experimental set-up

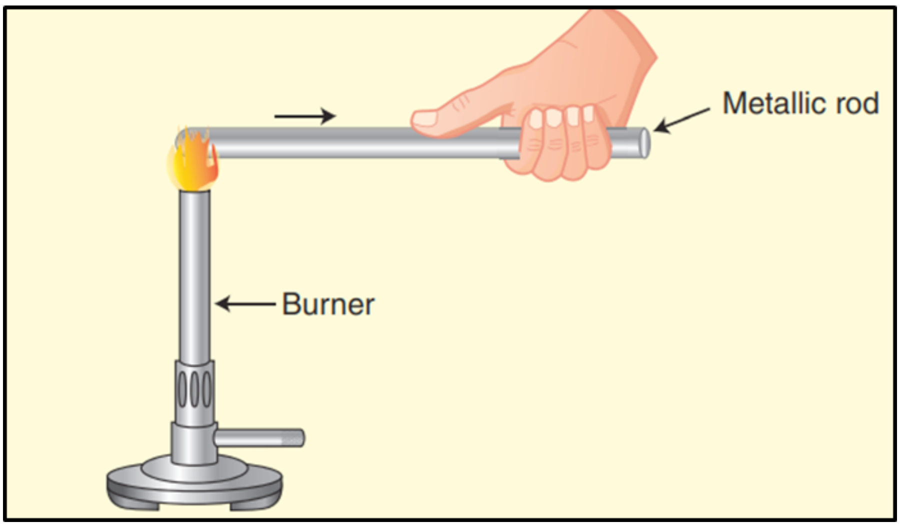


Figure 3.1. Conduction of heat in metals

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 experimental setup above;
3. Repeat the experiment by using sulphur rod instead of iron rod.

What do you feel about 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

In the experiment one iron rod conducts heat because it is a metal, while for the second experiment sulphur rod does not conduct heat because it is a 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

Experimental set-up

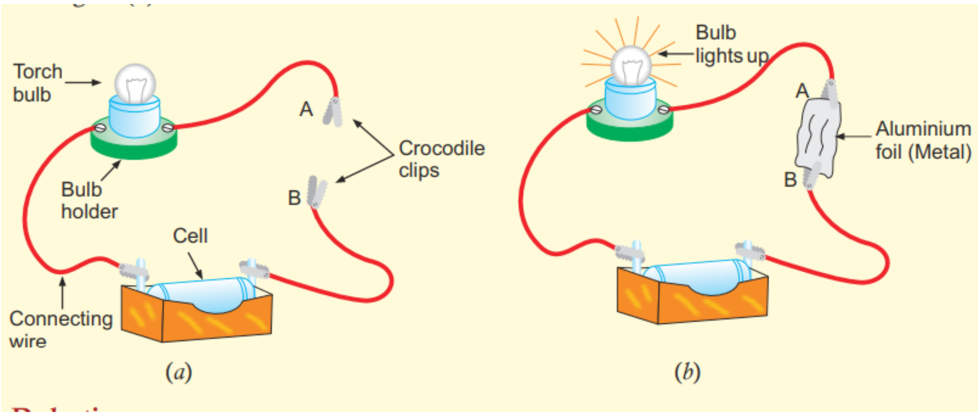


Figure 3.2. Electrical conductivity of metals and non-metals

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

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

UNIT: 6

TRENDS IN CHEMICAL PROPERTIES OF GROUP 1 ELEMENTS AND THEIR COMPOUNDS

EXPERIMENT 6.1: COMPARISON OF THE REACTIVITY OF SODIUM AND POTASSIUM WITH WATER

Rationale

Knowledge on the reaction of group 1 elements with water, provides skills on how to handle and store them. Comparing the reactivity of sodium and potassium will help to explain the position of group 1 elements in the periodic table with respect to their reactivity.

Objective

Learners will be able to compare the reactivity of sodium and potassium with water.

Required materials

Apparatus

- Beakers
- Knife
- Pair of tongs

Chemicals

- Water
- Sodium metal
- Potassium metal

Experimental set-up

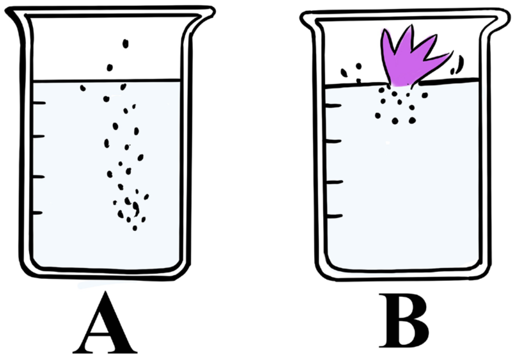


Figure 6.1: Comparison of the reactivity of sodium and potassium

Procedure

1. Take two Erlenmeyer flasks, and label them as A and B. Fill each of them with water up to three-fourth.
2. Cut a small piece of sodium metal with a knife and put it in beaker A using a pair of tongs. What do you observe?

   Answer*: Sodium reacts violently forming a ball-like structure that moves around on the surface of water with a hissing sound and evolution of a colourless gas.*
3. Cut a small piece of potassium metal with a knife and put it in a beaker B using a pair of tongs. What do you observe?

   Answer*: Potassium reacts much more violently and a colourless gas which ignites on contact with air is produced.*
4. Test the obtained solutions (in beaker A and B) with red and blue litmus papers and record your observations.

   Answer*: Red litmus paper turns blue, but the blue litmus paper does not change.*
5. Test the gas evolved by burning it with a lighting splint. What do you observe?

   Answer*: The gas burns with a pop sound*.

Interpretation of results and conclusion

Guiding questions

1. What is the identity of the products formed when group 1 elements react with water?
2. Write the chemical equations of the reactions which take place.

Answer to guiding question

Sodium and potassium react with water to form an alkaline solution which turns red litmus paper blue and the hydrogen gas which burns with a pop sound.

Potassium reacts more violently with water than sodium.

The equations of reactions

$$2\,Na_{(s)} + 2\,H_2O_{(l)} \longrightarrow 2\,NaOH_{(aq)} + H_{2(g)}$$

$$2\,K_{(s)} + 2\,H_2O_{(l)} \longrightarrow 2\,KOH_{(aq)} + H_{2(g)}$$

Alkali metals (M) react with water to form hydroxides (MOH) and hydrogen gas.

$$2\,M_{(s)} + 2\,H_2O_{(l)} \longrightarrow 2\,MOH_{(aq)} + H_{2(g)}$$

where M represents any group 1 metal and MOH represents the corresponding hydroxide.

The experiment shows that the reactivity of Group 1 metals increases as we move down the group because the electro-positivity increases.

Evaluation

1. State the expected observations:

   a) when sodium metal reacts with water

   b) when potassium metal reacts with water

   ***Answer**: See interpretation*

2. Compare the reactivity of sodium with that of potassium.

   ***Answer:** See interpretation*

3. Provide the balanced chemical equations of the following reactions, and mention the state symbols of the reactants and products

   a) sodium reacts with water

   b) potassium reacts with water

   Answer: *See interpretation*

4. Sodium and potassium react with water to give solutions. By introducing red and blue litmus papers to each of the obtained solutions, explain the observed colour changes.

   Answer: *See interpretation*

5. a) Which gas evolved when both sodium and potassium are reacted with water?

   b) Explain a chemical test to identify the gas produced in 5.a)

   Answer*: See interpretation*

6. Write the balanced chemical equation between Caesium, Rubidium, and Lithium with water

   Answer:

$$2Cs(s) + 2H_2O(l) \longrightarrow 2CsOH(aq) + H_2(g)$$

$$2Rb_{(s)} + 2H_2O_{(l)} \longrightarrow 2RbOH_{(aq)} + H_{2\ (g)}$$

$$2Li_{(s)} + 2H_2O_{(l)} \longrightarrow 2LiOH_{(aq)} + H_{2\ (g)}$$

EXPERIMENT 6.2: DEMONSTRATION OF ALKALINE CHARACTER OF SODIUM OXIDE

Rationale

Sodium oxide is a white solid which readily reacts with water to produce sodium hydroxide, a strong alkaline solution. It is mainly used in the manufacture of glasses and ceramics. This experiment aims to demonstrate the alkaline nature of group 1 oxides.

Objective

Learners will be able to demonstrate the alkaline character of sodium oxide.

Required materials

Apparatus

- Beaker
- Pair of tongs
- Stirring rod

Chemicals

- Water
- Sodium oxide
- Red and blue litmus paper

Procedure

1. Prepare a clean beaker, pour in 20 mL of distilled water.
2. Add 3 spatulas full of sodium oxide and mix with a stirring rod to dissolve the oxide. What do you observe?

   Answer: *Sodium oxide dissolves completely in water.*
3. Test the obtained solution with red and blue litmus papers. What do you observe?

   Answer: *Red litmus paper turns blue and blue litmus paper doesn't change*

Interpretation of results and conclusion

Guiding questions

1. What happens when sodium oxide is added to water?
2. The product obtained in question 1 is tested using blue and red litmus papers. Explain the observed changes.
3. a) Write the chemical equation of the reaction occurring when sodium oxide is dissolved in water by indicating the state symbols.

   b) State and explain the nature of the solution obtained in 3a).

Answer to guiding question

Sodium oxide dissolves in water. Aqueous solution obtained turns red litmus paper blue. Aqueous solution of sodium oxide is alkaline. Therefore, sodium oxide is a basic oxide in nature.

The equation of reaction is:

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

Oxides formed by group 1 metals readily react with water giving strong alkaline solutions.

$$M_2O(s) + H_2O(l) \longrightarrow 2MOH(aq)$$

Evaluation

Write a balanced chemical equation of each of the following oxides with water

a) Lithium oxide

Answer: $Li_2O(s) + H_2O(l) \longrightarrow 2LiOH(aq)$

b) Potassium oxide

Answer: $K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$

c) Rubidium oxide

Answer: $Rb_2O(s) + H_2O(l) \longrightarrow 2RbOH(aq)$

EXPERIMENT 6.3: DEMONSTRATION OF ALKALINE CHARACTER OF GROUP 1 HYDROXIDES

Rationale

Hydroxides of group 1 elements play important roles in our daily life. Sodium hydroxide and potassium hydroxide are used in soaps and detergent making. They are also applied in the manufacture of papers, textiles industries, and as drain cleaners. This experiment will help learners to handle the strong alkaline solutions with precautions due to their corrosive nature.

Objective

Learners will be able to demonstrate the alkaline character of group 1 hydroxides such as sodium hydroxide or potassium hydroxide.

Required materials

Apparatus

- Beakers
- Stirring rod

Chemicals

- Sodium hydroxide/ Potassium hydroxide
- Distilled water
- Phenolphthalein

Procedure

1. Avail two clean beakers and label them as A and B
2. Prepare a dilute solution of NaOH by dissolving about 1g in 25mL of water into beaker A. What do you observe?

   Answer: *Sodium hydroxide dissolved and the beaker got hot. This means that the process is exothermic*
3. Put about 5mL of distilled water into beaker B
4. Add about 2 drops of phenolphthalein indicator in each of the above beakers (A and B). What do you observe?

   Answer: *In beaker A, phenolphthalein indicator turns the solution pink whereas in beaker B there is no observable change.*

Interpretation of results and conclusion

Guiding question

1. Why does phenolphthalein turn pink in sodium hydroxide solution and remain unchanged in distilled water?
2. Write down the chemical equation of the dissociation of sodium hydroxide in water

Answer to guiding questions

The colour of phenolphthalein indicator turns pink in beaker A because the beaker contains an alkaline solution. Sodium hydroxide solution is an alkaline solution. In beaker B the colour of the indicator does not change because distilled water is neutral.

Equation of the reaction

$$NaOH(s) \longrightarrow Na^{+}(aq) + OH^{-}(aq)$$

All group 1 hydroxides dissolve in water to form a strong alkaline solution.

$$MOH(s) \longrightarrow M^{+}(aq) + OH^{-}(aq)$$

Evaluation

Carry out the same experiment by replacing sodium hydroxide with potassium hydroxide and interpret the obtained results

EXPERIMENT 6.4: ACTION OF HEAT ON GROUP 1 CARBONATES AND IDENTIFICATION OF THE FORMED PRODUCTS

Rationale

Carbonates of group 1 elements behave differently in the presence of heat. Lithium carbonate decomposes on heat while other carbonates do not. Decomposition of lithium carbonate leads to lithium oxide and carbon dioxide. Lithium oxide is an important compound mostly used in the production of ceramics and batteries while carbon dioxide is used in fire extinguishers, in making carbonated drinks and as raw material for photosynthesis. This experiment demonstrates the behaviour of group 1 carbonates when heated.

Objective

Learners will be able to explain the action of heat on Group 1 carbonates and identify products formed.

Required materials

Apparatus

- Glass test tubes
- Retort stand and accessories
- Bunsen burner
- Spatula
- Match box
- Test tube holders
- Delivery tube

Chemicals

- Lithium carbonate
- Sodium carbonate
- Potassium carbonate
- Lime water

Caution

Lithium oxide should be handled with care as it causes irritation, severe burns in contact with the skin and blindness in contact with eyes. For safety, remember to wear safety glasses (Goggles).

Use small amounts (not more than 2 spatulas end full) of solid substances during decomposition to avoid large amounts of toxic gases.

Experimental set-up

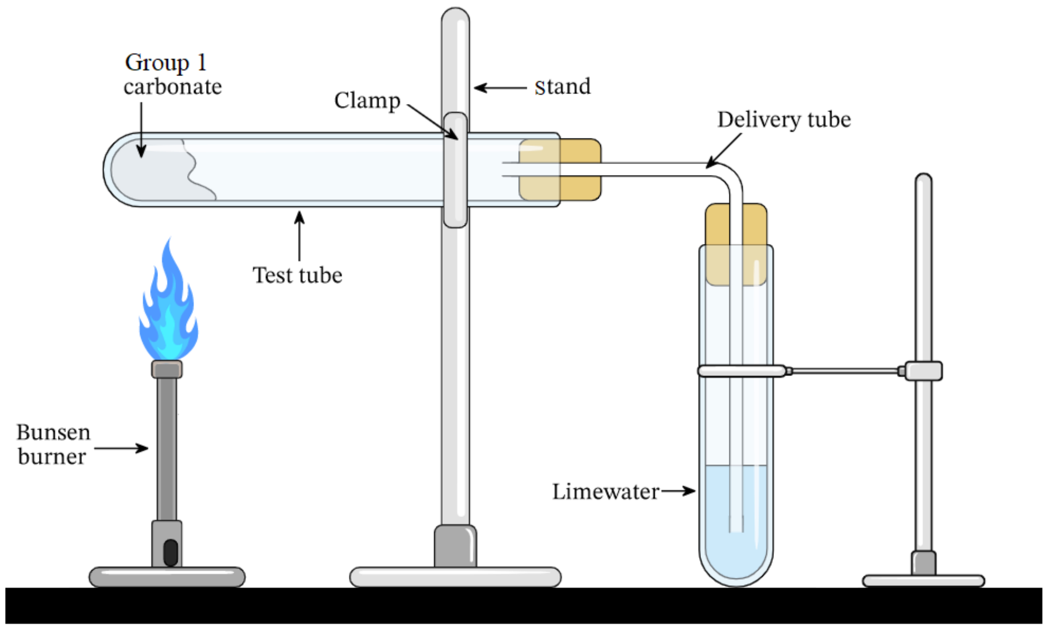


Figure 6.2: Thermal decomposition of group 1 carbonates

Procedure

Step I:

1. Take a spatula endful of lithium carbonate into a borosilicate test tube
2. Set the apparatus as shown in figure 6.2
3. Light a gas burner and heat lithium carbonate strongly until there is no further change
4. Test the gas evolved with limewater into another glass test tube as shown in the figure 6.2 Note your observation?

   Answer: *Lithium carbonate decomposes on heat flame to release a colourless gas which turns lime water milky.*

Step II:

5. Repeat the procedure in step I using potassium carbonate and sodium carbonate separately, and write down your observation.

   Answer: *Sodium and potassium carbonates do not decompose on heat flame to release a colourless gas which turns lime water milky.*

Data recording /observations

Compound	Observation	Inference
$Li_2CO_3(s)$	lime water turns milky	$Li_2CO_3(s)$ decomposes and releases a colourless acidic gas which turns lime water milky
$Na_2CO_3(s)$	No observable change	$Na_2CO_3(s)$ does not decompose easily
$K_2CO_3(s)$	No observable change	$K_2CO_3(s)$ does not decompose easily

Interpretation of results and conclusion

Guiding questions

1. How do group 1 carbonates behave when heated over a Bunsen burner flame?
2. Write the chemical reaction of the thermal decomposition of lithium carbonate.

Answer to guiding questions

In Group 1 carbonates, when heated on a Bunsen burner flame, only lithium carbonate decomposes to lithium oxide and carbon dioxide.

$$Li_2CO_3(s) \xrightarrow{\text{heat}} Li_2O(s) + CO_2(g)$$

Note: Group I carbonates are resistant to heat except lithium carbonate. This means that lithium carbonate is thermally less stable because the carbonate ion is more polarized by a small singly charged positive ion. The smaller the positive ion, the higher the charge density, and the greater the effect on the carbonate ion. This is the reason why lithium carbonate decomposes easily due to its low ionic character. The decomposition temperatures increase down the group.

Evaluation

1. Two carbonates X and Y of group 1 elements are heated strongly in a borosilicate test tube until there are no further changes. The carbonate X resists heat while the carbonate Y releases on heating a colourless gas which turns milky a colourless solution of lime water.

   a) Which carbonate among X and Y does correspond to

   i) Lithium carbonate and explain.

   ii) Sodium carbonate and explain.

   Answer: *See interpretation*

   b) i) Write the chemical equation of thermal decomposition of carbonate Y

   ii) Give the name and the formula of the gas released (W) by the thermal decomposition of carbonate Y.

   iii) Explain the reason why the released gas turns a colourless solution of lime water milky.

iv) Write the chemical equation of reaction between the released gas W and lime water.

Answer: *See interpretation*

c) The carbonate Y is said to be less stable than the carbonate X. Explain.

Answer: *See interpretation*

2. Which carbonate is more stable among the carbonates of group 1 element? Explain your answer

   Answer: *The more stable carbonate among group one carbonate is that of Caesium. This is because it has less polarizing power due to its large size.*

3. At high temperature, write a balanced chemical equation of the thermal decomposition of the following carbonates:

   a) Sodium carbonate

   Answer: $Na_2CO_3(s) \longrightarrow Na_2O(s) + CO_2(g)$

   b) Potassium carbonate

   Answer: $K_2CO_3(s) \longrightarrow K_2O(s) + CO_2(g)$

EXPERIMENT 6.5: ACTION OF HEAT ON GROUP 1 NITRATES

Rationale

The action of heat on group 1 nitrates highlights the thermal stability of them. The less stable lithium nitrate decomposes easily to its corresponding oxide, oxygen and nitrogen dioxide, while the remaining more stable nitrates decompose to their corresponding nitrites and oxygen. The experiment illustrates different behaviour between lithium nitrate and other group 1 nitrates when heated.

Objective

Learners will be able to explain the effect of heat on group 1 nitrates.

Required materials

Apparatus

- Test tubes
- Pair of tongs
- Wooden splint/ matchstick
- Bunsen burner/heat source
- Spatula
- Stopper with one hole
- Delivery tube
- Retort stand and accessories

Chemicals

- Lithium nitrate
- Sodium nitrate
- Potassium nitrate
- Blue and red litmus papers

Experimental set-up

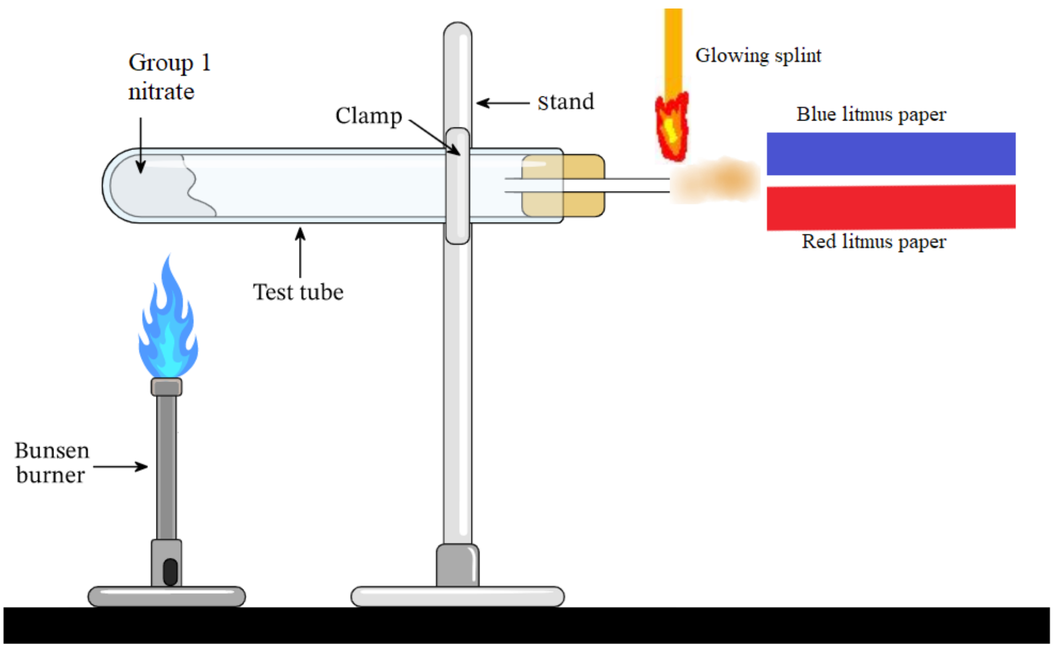


Figure 6.3: Thermal decomposition of group 1 nitrates

Procedure

1. Put two spatulas endfull of lithium nitrate into a borosilicate test tube. Heat it strongly until there is no further change and test the gases evolved if any with damp blue and red litmus papers and a glowing splint. Note your observation.

   Answer: *Lithium nitrate releases a brown gas which turns blue litmus paper red. The red litmus paper does not change. A glowing splint relights. A white residue is observed in the test tube.*

2. Repeat the procedure by using potassium nitrate and sodium nitrate and write your observation.

   Answer: *Potassium and sodium nitrates release a colourless gas that relights a glowing wooden splint and has no effect on litmus papers. A white residue is also observed in the test tube.*

Interpretation of results and conclusion

Guiding questions

1. What are the products formed during thermal decomposition of lithium nitrate?
2. From question 1, a) which gas does turn blue litmus paper red?

   b) Which gas relights a glowing splint?
3. What conclusion can be drawn from the action of heat on sodium nitrate and potassium nitrate?

Answers to guiding questions

The thermal decomposition of lithium nitrates produces lithium oxide (white residue), nitrogen dioxide (brown gas) and oxygen (colourless).

The chemical equation for the decomposition of $LiNO_3$ is

$4LiNO_3(s) \rightarrow 2Li_2O(s) + NO_2(g) + O_2(g)$

The produced nitrogen dioxide turns blue litmus paper to red. Thus, NO_2 is an acidic gas while oxygen relights a glowing splint.

Sodium and potassium nitrates decompose at higher temperatures than lithium nitrate and they break down when heated to give off oxygen as the only gas. The solid residue formed is sodium or potassium nitrite.

Except lithium nitrate, on heating group 1 nitrates decompose to nitrites and oxygen.

$$2NaNO_{3(s)} \longrightarrow 2NaNO_{2(s)} + O_{2(g)}$$

The general chemical equation for the decomposition reaction of group1 metal nitrates is

$$2MNO_{3(s)} \longrightarrow 2MNO_{2(s)} + O_{2(g)}$$

where M represents a group 1 metal other than Li.

Evaluation

1. i) Write the chemical equation of thermal decomposition of potassium nitrate and name the formed products

   Answer: *See general equation*

   ii) Explain the chemical test for the produced gas in 1.i)

   Answer: *See interpretation*

2. i) Write the chemical equation of thermal decomposition of lithium nitrate and name the formed products

   ii) Explain a chemical test that could be used to distinguish between the two formed gases in 2.i)

   Answer: *See interpretation*

EXPERIMENT 6.6: SOLUBILITY OF GROUP I COMPOUNDS

Rationale

In general, the solubility of group 1 compounds makes them easily applied in different fields such as agriculture, environment, industries, and medicines. Knowledge in the trends of group 1 compounds solubility help to predict their solubility order during solution preparations. This is helpful and applied in the process of solvent extraction and water purification. The experiment deals with the trend in solubility of group 1 compounds.

Objective

Learners will be able to explain the solubility of group 1 compounds.

Required materials

Apparatus

- Beakers
- Test tubes
- Stirring rod
- Spatulas

Chemicals

- Distilled water
- Potassium hydroxide
- Sodium carbonate
- Sodium sulphate
- Potassium sulphate
- Lithium hydroxide
- Rubidium hydroxide
- Lithium carbonate
- Potassium carbonate
- Sodium hydroxide
- Phenolphthalein

Procedure

Experiment I

1. Prepare 4 clean beakers and label them as A, B, C, and D
2. Pour 50 mL of water into each beaker and add a few drops of phenolphthalein indicator.
3. Add 2g of lithium hydroxide in a beaker A, 2g of sodium hydroxide in beaker B, 2g of potassium hydroxide in a beaker C, and 2g of Rubidium hydroxide in beaker D and stir with a stirring road. What do you observe?

   ***Answer**: Potassium hydroxide dissolves in water easily and phenolphthalein turns progressively pink as the solution is being formed.*

Experiment II

4. Repeat experiment I using metal carbonates of group 1 elements. Interpret your observation.
5. Repeat experiment I using metal sulphates of group 1 elements. Interpret your observation.

Note: Do not add phenolphthalein in experiments involving carbonates and sulphates.

***Answer:** Hydroxides, Carbonates and Sulphates of group 1 are soluble.*

Interpretation of results and conclusion

Guiding question

How is the trend of solubility of group 1 compounds?

Answer to guiding question

Solubility is the maximum amount of a substance that dissolves in a given amount of solvent at a specified temperature.

Group 1 carbonates are very soluble. The least soluble of group 1 carbonate is lithium carbonate. The solubility of carbonates increases as you go down the group.

Hydroxides of group 1 are even more soluble. Lithium hydroxide is the least soluble. Solubility of hydroxides increases as you go down the group.

The solubility of sulphates behaves in different way from carbonates and hydroxides. The solubility of group 1 sulphates decreases as you go down the group.

Some lithium compounds are less soluble in water due to their covalent character caused by the high polarizing power of Li cation. For example, LiI is more soluble in CCl_4, a non-polar solvent, than in H_2O which is polar.

Evaluation

Arrange the following compounds in ascending order of their solubility and explain

a) RbOH, NaOH, LiOH, CsOH

b) Cs_2CO_3, Li_2CO_3, Na_2CO_3, K_2CO_3

***Answer**: See interpretation*

EXPERIMENT 6.7: IDENTIFICATION OF GROUP 1 CATIONS

Rationale

Flame test is one of the analytical tests used to identify certain metal ions. Group 1 metal ions are identified easily as they generate different colours in the flame test. The colours of flame of various ions find their use in fireworks.

Objective

Learners will be able to identify different cations like Li^+, Na^+ and K^+ in a given solution using a flame test.

Required materials

Apparatus

- Bunsen burner
- Test tubes
- Magnesia rod/ Nichrome wire

Chemicals

- Lithium carbonate or lithium chloride
- Potassium sulphate or Potassium chloride
- Sodium sulphate or sodium chloride

Experimental set-up

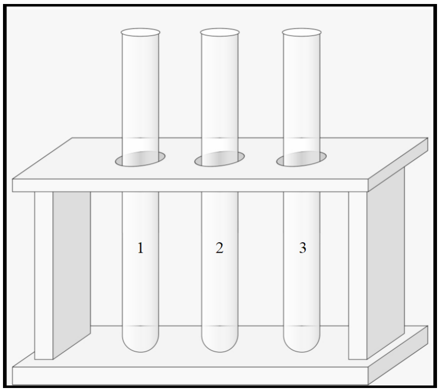


Figure 6.5: Set up for flame tests of alkali metals

Procedure

1. Dissolve a small amount of lithium carbonate, sodium chloride and potassium sulphate in test tube 1, 2, and 3 containing water respectively.
2. Wash the magnesia rod/nichrome wire to remove impurities
3. Dip the magnesia rod/nichrome wire in test tube 1 to get some drops of the solution of lithium carbonate.
4. Light the Bunsen burner
5. Gently wave the magnesia rod/nichrome wire on a non-luminous Bunsen burner flame.
6. Repeat steps 2 to 5 for the other two solutions (sodium chloride and potassium sulphate). What do you observe in each case?

   ***Answer**: Lithium carbonate produces a crimson red flame; sodium chloride produces a yellow flame and potassium sulphate produces a lilac/light purple flame*.

Interpretation of results and conclusion

Guiding question

What happens to the alkali metal cations when heated on a non-luminous Bunsen burner flame?

Answer to guiding question

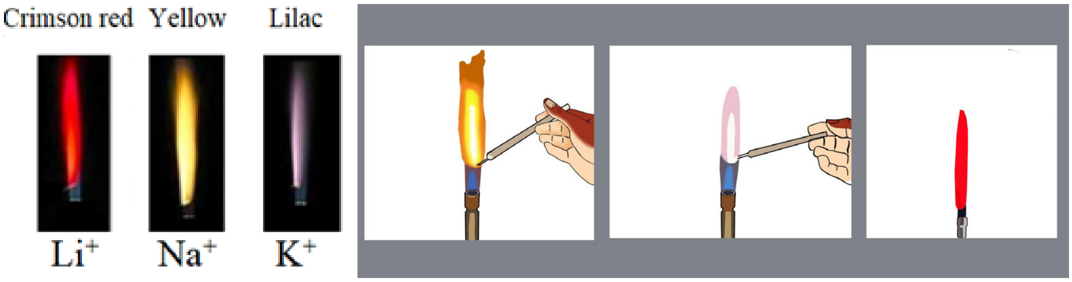


***Figure 6.6: Flame colours of Na^+, K^+, and Li^+** respectively*

Note on sources of errors:

The expected flame colour may not be clearly seen due to the impurities present on the magnesia rod/nichrome wire. So, if you are using one make sure to wash it before testing another solution with distilled water or aqueous HCl.

Evaluation

1. a) A flame test has been performed and the following observations were recorded on different solutions as shown in the table below. Complete the blank spaces.

Solution	Flame test observation	Expected ions present in the solution
A	*Light purple colour*	
B		lithium ions present
Sodium carbonate		

   b) Explain the origin of colours when flame test is conducted on solutions containing ions

2. Describe a chemical test used to distinguish between each pair of the following salt solutions below and state the expected results.

   a) Solutions of CsCl and KCl

   b) Solutions of NaCl and LiCl

   ***Answer**: See interpretation*

UNIT: 7

TRENDS IN CHEMICAL PROPERTIES OF GROUP 2 ELEMENTS AND THEIR COMPOUNDS

EXPERIMENT 7.1: COMPARISON OF THE REACTIONS OF MAGNESIUM AND CALCIUM WITH WATER AND DILUTE HYDROCHLORIC ACID

Rationale

Magnesium, the second element of group 2, plays many crucial roles in the body, such as supporting muscle and nerve function and energy production. Calcium is a mineral most often associated with healthy bones and teeth. Comparing the reactivity of magnesium and calcium will help to predict and explain the position of group 2 elements in the periodic table with respect to their reactivity.

Objective

Learners will be able to compare the reactivity of magnesium and calcium with water and dilute acids.

Required materials

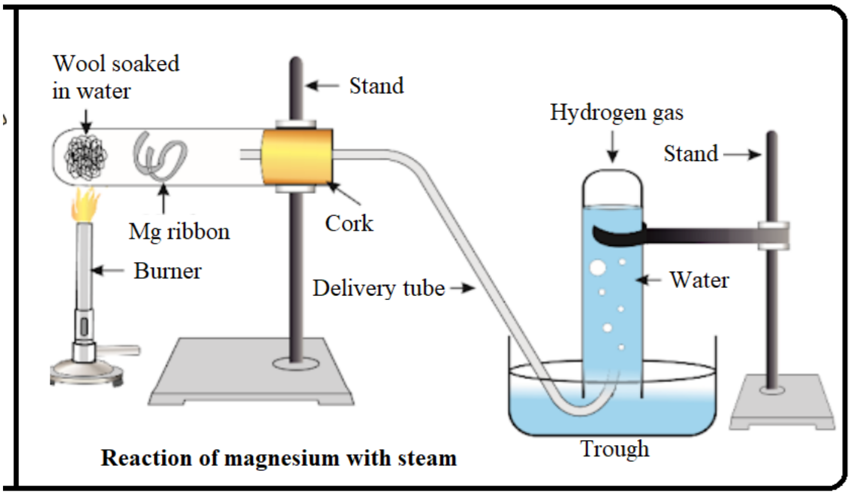
Apparatus

- Test tubes
- Beakers
- Rubber stoppers with two holes
- Delivery tubes
- Bunsen burner
- Match box, wooden splint
- Dropping funnel
- Trough

Chemicals

- Magnesium ribbon
- Calcium metal
- 1M hydrochloric acid
- Blue and red litmus papers
- Water
- NaOH

Experimental set-up



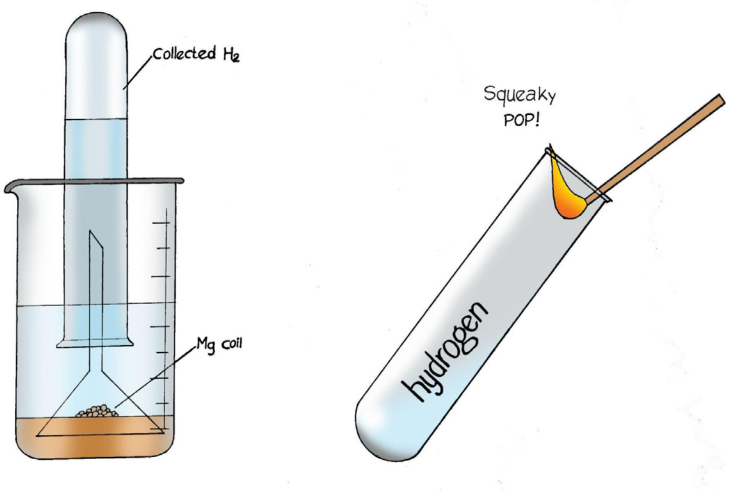


Figure 7.1: Reactions of magnesium and calcium with water

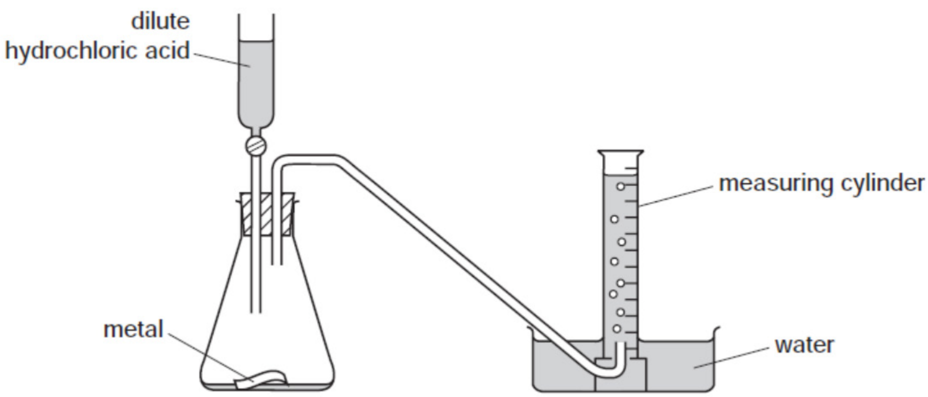


Figure 7.2: Reactions of magnesium or calcium with dilute hydrochloric acid

Caution:

- Wear splash-proof goggles and tell students to keep a distance of about 2 meters.
- Never look directly at a burning magnesium ribbon.
- Do not use magnesium powder.
- Do not add calcium to hot water.

Procedure

I. **Reactions of magnesium and calcium with water**

1. Put about 100mL of cold water in clean beaker
2. Add about 0.5g of calcium. What is your observation?

   Answer*: Liberation of a colourless gas and the reaction produces heat.*
3. Invert a filter funnel in the reacting beaker and collect the gas evolved in a test tube as indicated in Figure 7.1.
4. Put a burning splint over the opening of the test tube. What do you observe?

   Answer*: The colourless gas burns with a pop sound.*
5. Repeat steps 1 and 2 using magnesium ribbon.

   Observation**:** No *observable change*
6. Repeat step 5 using magnesium ribbon and hot water (around 80°C).

   Observation: *A colourless gas that burns with a pop sound is liberated.*
7. Considering the results of step 6, test the resulting solutions with red and blue litmus paper and note your observations.

   Answer: *Red litmus paper turns blue while blue litmus paper does not change.*

II. **Reactions of magnesium and calcium with dilute hydrochloric acid**

1. Put about 10 mL of 1M HCl in a clean boiling tube.
2. Add about 2-5 cm length of magnesium ribbon (Mg should be in excess). What do you observe?

   Answer*: A colourless gas is evolved*

3. Put a burning splint over the opening of the boiling tube. What do you observe?

   Answer: *The evolved gas burns with a pop sound.*

4. Divide the resultant solution into two portions.
5. Test the first portion with wet litmus papers (blue and red). What do you observe?

   Answer: *Red litmus paper turns blue while the blue one remains unchanged.*

6. To the second portion, add NaOH solution drop wise until in excess. Note your observation?

   Answer: *A white precipitate which is insoluble in excess of NaOH is formed.*

Data recording/ observations

Reaction	Observations	Inference
Reaction of calcium and magnesium with cold water	There is liberation of a colourless gas that burns with a pop sound when calcium is used. The reaction produces heat.	The generated gas is hydrogen……….
	Magnesium does not react with cold water	N/A
Reaction of magnesium with steam.	There is liberation of a colourless gas that burns with a pop sound.	
Reaction of magnesium with dilute acid	There is liberation of a colourless gas that burns with a pop sound	

Interpretation of results and conclusion

Guiding questions

What are the products of the reaction of:

a) magnesium with cold, hot water and steam?

b) calcium with cold water?

c) magnesium with dilute hydrochloric acid?

d) calcium with dilute hydrochloric acid?

Answer to guiding questions

Calcium readily reacts with cold water to give calcium hydroxide and hydrogen. Magnesium reacts very slowly with cold water. It reacts quickly with hot water to give magnesium hydroxide and water.

The chemical equations of the reactions that take place are:

$$Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$$

$$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

Note: Magnesium reacts with steam to produce magnesium oxide and hydrogen gas.

$$Mg(s) + 2H_2O_{steam} \longrightarrow MgO(s) + H_{2(g)}$$

Mg reacts with 1M HCl to produce hydrogen gas and Mg^{2+} ions which react with $NaOH_{(aq)}$ to give a white precipitate of magnesium hydroxide which is insoluble in excess of NaOH.

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

$$Mg^{2+}(aq) + 2OH^-(aq) \longrightarrow \underset{\text{White precipitate}}{Mg(OH)_2(s)}$$

$$MgCl_2(aq) + 2NaOH(aq) \longrightarrow Mg(OH)_2(s) + 2NaCl(aq)$$

Calcium is more reactive in water and in acids than magnesium and the reactivity of group 2 elements increases as you go down the group.

Evaluation

1. a) Calcium is more reactive than magnesium. Justify this assertion by referring to the reactivity of those metals with cold water and the steam of water.

   Answer: *See interpretation*

   b) Write the chemical equations for the following reactions. In each reaction state your observation

   i) magnesium and cold water

   ii) magnesium and steam of water

   iii) calcium and cold water

   Answer: *See interpretation*

   c) Explain a chemical test to identify the evolved gas in b) and describe your observations.

   Answer: *See interpretation*

2. Write the chemical equations of reactions between Mg and dilute hydrochloric acid and name the formed products.

   Answer: *See interpretation*

3. Explain why magnesium does not readily react with cold water.

   Answer: Magnesium does not readily react with cold water because in the air it is covered with a thin layer of magnesium hydroxide which prevents it from reacting with water.

4. Explain why calcium and sodium react violently with acids?

   Answer: *Calcium and sodium are stronger reducing agents than hydrogen and therefore they easily displace it from acids.*

EXPERIMENT 7.2: ALKALINE CHARACTER OF GROUP 2 OXIDES AND HYDROXIDES

Rationale

Calcium oxide is widely used in making cements, mortars, glasses and porcelain; in purifying sugar; in preparing bleaching powder, in water softeners. This compound is used as a pH modifier due to its basicity specifically by being used to reduce soil acidity. The production of many plastics involves the use of calcium hydroxide as an ingredient. Magnesium oxide is used as an antacid to relieve heartburn or acid indigestion due to its basicity. This experiment demonstrates the basic character of oxides and hydroxides of group 2 elements.

Objective

Learners will be able to demonstrate the alkaline character of group 2 oxides and hydroxides.

Required materials

Apparatus

- Beaker 250mL
- Stirring rod
- Dropper
- Spatula
- Thermometer

Chemicals

- Calcium oxide
- Distilled water
- Phenolphthalein / red litmus paper

Procedure

1. Place 2g of calcium oxide in a beaker.
2. Add slowly 100 mL of distilled water.
3. Shake the mixture well or mix with a stirring rod. record your observation.

   Answer: *Calcium oxide slightly dissolves in water to form a suspension and the mixture heats up.*

4. Leave the mixture in the beaker for some time and note your observation.

   Answer: A clear solution is formed and the solid particles settle down the solution

5. Add 2 drops of phenolphthalein or test the mixture with a red litmus paper. Write your observation.

   Answer: *Red litmus paper turns blue or the solution turns pink in presence of phenolphthalein.*

Interpretation of results and conclusion

Guiding questions

a) What is the name and chemical formula of the obtained product when calcium oxide is added into water?

b) Is calcium oxide an acidic, basic or neutral oxide? Explain.

Answer to guiding questions

When a small amount of calcium oxide is dissolved in water, it forms a solution that turns red litmus paper blue showing that its solution is alkaline. Calcium oxide is a basic oxide.

The chemical equation of reaction between calcium oxide and water is:

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$$

When calcium hydroxide solution is left undisturbed for some time, a small portion of it dissolves forming an alkaline solution commonly called lime water.

The chemical equation of reaction:

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$$

$$Ca(OH)_2(aq) \longrightarrow Ca^{2+}(aq) + 2OH^{-}(aq)$$

Note: Except those of beryllium which are amphoteric, group II oxides and hydroxides are basic, and the oxides react with water to produce corresponding hydroxides. The group II hydroxides become more soluble in water as you go down the group.

$$MO(s) + H_2O(l) \longrightarrow M(OH)_2(s)$$

$$M(OH)_2(s) + H_2O(l) \longrightarrow M^{2+}(aq) + 2OH^{-}(aq) \text{ alkaline solution}$$

Evaluation

1. What does a basic oxide mean?

   Answer: *Basic oxide is an oxide which forms an alkaline solution when dissolved in water*

2. When limestone is heated at high temperature,

   a) It forms calcium oxide or quicklime.

   b) When water is added to quicklime, slaked lime (ishwagara) is formed.

   Write the chemical equations involved in a) and b).

   Answer: *See interpretation*

EXPERIMENT 7.3: ACTION OF HEAT ON CARBONATES OF GROUP 2 ELEMENTS

Rationale

All group 2 carbonates undergo thermal decomposition to metal oxide and carbon dioxide gas. They act as raw materials in different industrial processes such as drug development, glass making, soap and detergent production, paper industry and water softener. This activity deals with carbonates of group 2 elements.

Objective

Learners will be able to explain the effect of heat on carbonates of group 2 elements.

Required materials

Apparatus

- Borosilicate test tube
- Bunsen burner
- Two corks with one hole
- Delivery tube

Chemicals

– Calcium carbonate
– Lime water
– Retort stand

Experimental set-up

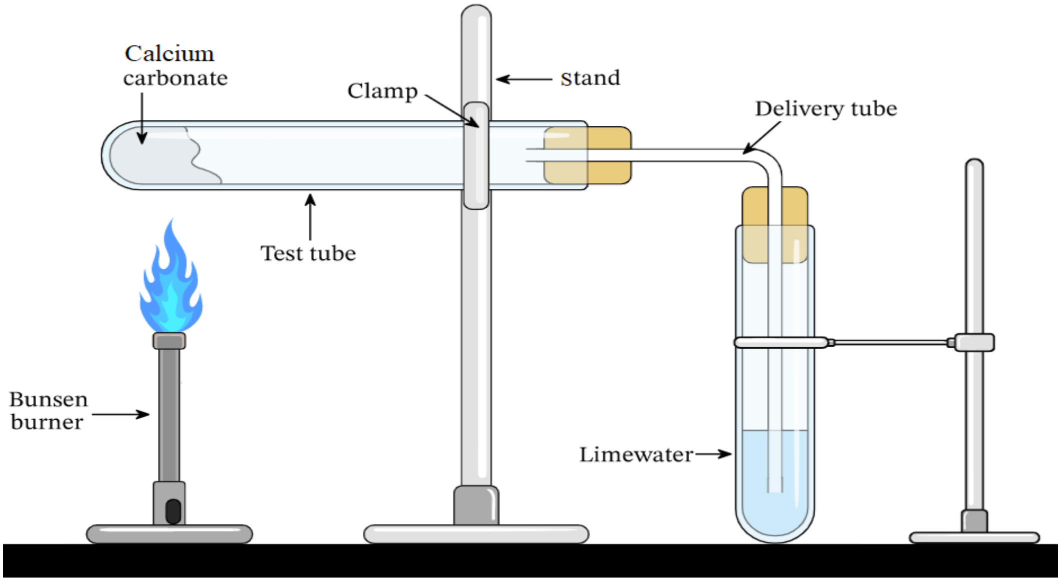


Figure 7.3: Decomposition of calcium carbonate

Procedure

1. Put 2 g of calcium carbonate ($CaCO_3$) in a borosilicate test tube and 5 mL of lime water into another test tube
2. Set up the apparatuses as shown in the figure 7.3 and start heating gently. What do you observe?

   Answer: *Calcium carbonate decomposes, and the gas evolved is observed in the test tube*
3. Heat strongly and state your observation when the evolved gas reaches in the test tube containing lime water.

   Answer: *The evolved gas turns lime water milky, and a whitish solid product is left in the borosilicate test tube.*

Interpretation of results and conclusion

Guiding question

What is the formed product when calcium carbonate undergoes thermal decomposition?

Answer to guiding question

When heated, calcium carbonate decomposes to calcium oxide (white) solid and carbon dioxide gas.

Chemical equation of the reaction:

$$CaCO_{3}(s) \xrightarrow{\text{heat}} CaO(s) + CO_{2}(g)$$

When carbon dioxide bubbles in lime water, a clear solution turns milky.

On heating, all carbonates of group 2 elements decompose to their corresponding metal oxide and carbon dioxide gas. During thermal decomposition of carbonates, the heat needed increases as we move down the group because as the cation size increases its polarizing power decreases which makes the carbonate more ionic with stronger electrostatic forces.

$$MCO_{3}(s) \xrightarrow{\text{heat}} MO(s) + CO_{2}(g)$$

Where M represents a metal of group 2 elements

Evaluation

1. Write the chemical equation of thermal decomposition of magnesium carbonate.

   Answer: $MgCO_3(s) \xrightarrow{\text{heat}} MgO(s) + CO_2(g)$

2. You have two carbonates A and B. When B is heated over a Bunsen burner flame decomposes, and a gas evolves while A does not decompose. Which one corresponds to sodium carbonate, and which one corresponds to calcium carbonate? Justify your answer.

   Answer: *B is calcium carbonate because a gas is evolved when it decomposes but sodium carbonate (A) does not decompose*

EXPERIMENT 7.4: ACTION OF HEAT ON NITRATES OF GROUP 2 ELEMENTS

Rationale

The nitrates of group 2 elements are mainly used as a fertilizer in agriculture. Calcium nitrate plays a big role in the maintenance of healthy cell walls in plants because of its higher solubility in the soil. It also helps the plants in absorbing other important nutrients in the soil. The experiment illustrates the behaviour of the nitrates of group 2 elements when subjected to heat.

Objective

Learners will be able to explain the effect of heat on nitrates of group II elements.

Required materials

Apparatus

- Borosilicate test tube
- Bunsen burner
- Wooden splint
- Match box
- Retort stand

Chemicals

- Magnesium nitrate
- Blue and red litmus paper

Experimental set-up

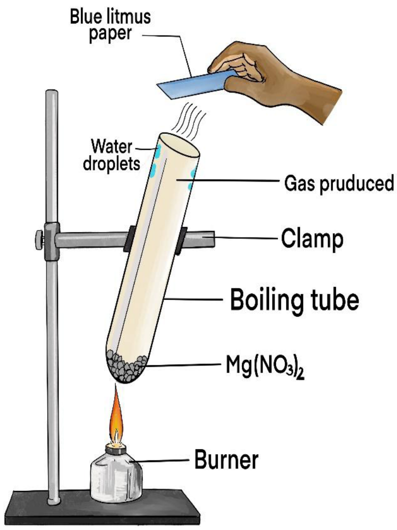


Figure 7.4: Decomposition of Magnesium nitrate

Procedure

1. Put 2 g of magnesium nitrate $Mg(NO_3)_2$ in a borosilicate test tube and heat gently. What do you observe?

   Answer: *Magnesium nitrate melts and decomposes. Water vapours are produced and condense on the wall of the test tube.*

2. Proceed with strong heating. Note your observation.

   Answer*: A brown gas is released, and a white product remains in the test tube.*

3. Insert a glowing wooden splint into the test tube. Write down your observations.

   Answer: *The glowing splint relights inside the test tube.*

4. Place a wet blue litmus paper near the opening of the test tube and test the released gas. Repeat this step with a wet red litmus paper. Record your observations.

   Answer: *The brown gas turns blue litmus paper red but the red one does not change.*

Interpretation of results and conclusion

Guiding question

What happens to magnesium nitrate when heated?

Answer to guiding question

Magnesium nitrate decomposes when heated. It produces magnesium oxide (the white product that remains in the test tube), nitrogen dioxide gas (the brown fumes) and oxygen gas (which relights a glowing wooden splint).

The chemical equation of the reaction:

$$2Mg(NO_3)_2(s) \xrightarrow{heat} 2MgO(s) + 4NO_2(g) + O_2(g)$$

All nitrates of group 2 elements decompose on heating to a metal oxide, nitrogen dioxide and oxygen. During thermal decomposition of nitrates, the heat needed increases as we move down the group because as the cation size increases its polarizing power decreases which makes the nitrates more ionic with stronger electrostatic forces.

$2M(NO_3)_{2(S)} \xrightarrow{heat} 2MO_{(s)} + 4NO_{2(g)} + O_{2(g)}$ where M represents a group II metal

Evaluation

1. Write the chemical equation of the thermal decomposition of calcium nitrate

   Answer: $2Ca(NO_3)_2$ Heat→ $2CaO(s)+4NO_2(g)+O_2(g)$

2. Two nitrates A and B are strongly heated. The nitrate B easily decomposes to give a white residue and a mixture of two gases whereas the nitrate A does not easily decompose. Which one do you expect to be $Mg(NO_3)_2$ and which one is $Ba(NO_3)_2$. Justify your answer.

   Answer: *B is magnesium nitrate and A is barium nitrate because magnesium ions have a higher polarizing power than barium ions.*

EXPERIMENT 7.5: SOLUBILITY OF GROUP 2 COMPOUNDS (HYDROXIDES AND SULPHATES)

Rationale

Group II metal hydroxides and sulphates become more and less soluble in water as you go down the group respectively. They are used in various aspects of life. For instance, calcium hydroxide is used in industrial settings such as sewage treatment, paper production, construction, and food processing. Magnesium hydroxide is used to treat occasional constipation in children and adults on a short-term basis and it is used to treat stomach acidity (milk of magnesia). Calcium sulphate provides a natural source of calcium and sulphur which can be directly assimilated by plants and are vital to fertilisation and healthy plant growth.

Objective

Learners will be able to explain the solubility of group II hydroxides and sulphates.

Required materials

Apparatus

- Test tubes
- Spatula
- Test tube rack
- Labels

Chemicals

- Distilled water
- Magnesium hydroxide
- Calcium hydroxide
- Magnesium sulphate
- Calcium sulphate
- Barium hydroxide
- Barium sulphate

Procedure

1. Prepare six test tubes and label them A, B, C, D, E, and F
2. Put 3 mL of water into each test tube
3. Add a half spatula of each alkaline earth metal (group 2 metals) compounds as shown below:
    - A= Magnesium hydroxide
    - B= Magnesium sulphate
    - C= Calcium hydroxide
    - D= Calcium sulphate
    - E= Barium hydroxide
    - F= Barium sulphate
4. Shake the test tubes to make aqueous solution. Record your observations in the following table.

Group II hydroxides compounds	Observations	Group II sulphates compounds	Observations
$Mg(OH)_2$	Slightly soluble	$MgSO_4$	Very soluble
$Ca(OH)_2$	Partially soluble	$CaSO_4$	Partially soluble
$Sr(OH)_2$	Soluble	$SrSO_4$	Insoluble
$Ba(OH)_2$	Very soluble	$BaSO_4$	Very insoluble

Interpretation of results and conclusion

Guiding question

How does the solubility of group II hydroxides and sulphates change in the group?

Answer to guiding question:

The solubility of group II hydroxides increases as you go down the group while that of group 2 sulphates decreases down the group.

Generally, group II compounds with small sized anions such as hydroxides and chlorides become more soluble in water as we move down the group. Their lattice energies decrease more than their hydration energies. Contrary, compounds with large sized anions such as sulphates and carbonates become less soluble downward as their hydration energies decrease more than lattice energies.

Evaluation

1. i) State your observation when:

    a) Crystals of barium hydroxide are put into a beaker containing distilled water

    b) The solution in a) is tested using phenolphthalein indicator.

    ii) What are your deductions about the nature of solution of barium hydroxide.

2. Write the ionic equation of the dissolution of barium hydroxide in water.

3. a) State what is observed when crystals of barium sulphate are put into a beaker containing distilled water

    b) Explain and compare the solubility of group II hydroxides vis-a-vis the solubility of group II sulphates.

    c) Explain why the solubility of group II hydroxides increases while that of sulphates decreases down the group?

4. Arrange the following compounds in ascending order of their solubility in water

    a) $Ba(OH)_2$

    b) $Mg(OH)_2$

    c) $Sr(OH)_2$

    d) $Ca(OH)_2$

    Answer: *$Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$*

    "<", "less soluble than"

EXPERIMENT 7.6: CHEMICAL TEST FOR CATIONS OF GROUP 2 ELEMENTS

They are two ways of testing group 2 elements: test using reactants and flame test.

a) Flame test

Rationale

Flame test is one of the analytical tests used to identify certain metal ions. Group 2 metal ions are identified easily as they generate different colours in the flame test. The colours of flame of various ions find their use in fireworks.

Objective

Learners will be able to practically distinguish magnesium, calcium and barium cations.

Required materials

Apparatus

- Spatulas
- Watch glass
- Bunsen burner
- Nichrome wire

Chemicals

- Distilled water
- Dilute Hydrochloric acid
- Magnesium chloride
- Calcium chloride
- Strontium chloride
- Barium chloride

Procedure

1. Pour dilute hydrochloric acid into a 50 ML beaker
2. Dip nichrome wire in the HCl to clean it.
3. Light the Bunsen burner
4. Heat the nichrome wire until blue flame is seen to clean it

5. Take calcium chloride solid from a watch glass using chrome wire
6. Heat it at the edge of the Bunsen burner flame
7. Repeat step 2 to 5 with magnesium chloride, strontium chloride and barium chloride

What do you observe?

Answer:

Ion present	Flame test colour
Magnesium, Mg^{2+}	Blinding white light
Calcium, Ca^{2+}	Brick red flame
Strontium, Sr^{2+}	Crimson red flame
Barium, Ba^{2+}	Pale green flame

Interpretation of results and conclusion

Guiding question

Describe the test of group 2 ions using the flame

Answer to guiding question

The flame test is used to determine visually the identity of an unknown metal ion based on the characteristic color the salt turns the flame of a Bunsen burner. The heat of the flame converts the metal ions into atoms which become excited and emit visible light.

The table above shows the flame test colours for $Mg^{2+,}$ Ca^{2+}, Sr^{2+} and Ba^{2+}.

Evaluation

1. What is the purpose of dipping the nichrome wire in hydrochloric acid before using it to test the given ion?
2. Which colour is produced when the following compounds are tested?
   - Calcium chloride
   - Magnesium chloride
   - Strontium chloride
   - Barium chloride

***Answer**: See interpretation*

b) **Chemical test for cations using chemicals**

Rationale

Group 2 compounds are all white solids at room temperature and pressure. Due to this, their physical appearance is not enough to distinguish them. Chemical methods are used to have reliable information about their true identities. Qualitative analysis of group 2 cations finds application in water analysis, soil composition determination, and in other industrial analysis. This experiment involves the use of analytes with selected reagents to give observable changes that we use to distinguish the analyzed samples.

Objective

Learners will be able to practically distinguish magnesium, calcium and barium cations.

Required materials

Apparatus

- Test tubes
- Spatulas
- Test tube rack
- Beakers
- Droppers

Chemicals

- Mg^{2+} solution
- Ca^{2+} solution
- Ba^{2+} solution
- Ammonia solution
- Sodium hydroxide solution
- Sodium sulphate solution

Procedure

1. Prepare 1M of each of the following solutions: ammonia, sodium hydroxide and sodium sulphate solutions.
2. Take three test tubes and label them as A, B and C.
3. Separately, dissolve some soluble salt of Mg, Ca and Ba compounds in test tubes A, B and C containing some amount of water respectively.

4. Divide each of the solutions obtained in step 3 into three separate test tubes and label them as A_1, A_2, A_3, B_1, B_2, B_3, and C_1, C_2, C_3 respectively.
5. Put test tubes A_1, B_1 and C_1 in a test tube rack and add aqueous sodium hydroxide dropwise until in excess. What do you observe?

   Answer: *In test tubes A_1 and B_1, a white precipitate is formed which is insoluble in excess sodium hydroxide whereas in test tube C_1 there is no observable change.*
6. Put test tubes A_2, B_2 and C_2 in a test tube rack and add aqueous ammonium hydroxide dropwise until in excess. What do you observe?

   Answer: *In test tube A_2 a white precipitate is formed which is insoluble in excess sodium hydroxide but in test tubes B_2 and C_2 there is no observable change.*
7. Put test tubes A_3, B_3 and C_3 in a test tube rack and add some drops of aqueous sodium sulphate. What do you observe?

   Answer: *In test tube A_3 there is no observable change. In test tube B_3 the solution becomes milky and in test tube C_3 a dense white precipitate is formed.*

Data recording

Addition of NaOH(aq)	Addition of NH_3(aq)	Addition of Na_2SO_4(aq)	Deduction
White precipitate, of $Mg(OH)_2$, insoluble in excess formed	White precipitate, of $Mg(OH)_2$, insoluble in excess formed	No change	Mg^{2+} present
White precipitate, of $Ca(OH)_2$, insoluble in excess formed	No change	Milky solution containing $CaSO_4$ formed	Ca^{2+} present
No change	No change	Dense white precipitate of $BaSO_4$ formed	Ba^{2+} present

Interpretation of results and conclusion

Guiding question

What is the chemical test used to identify the cations of group 2 elements?

Answer to guiding question

Group 2 compounds have different solubility, and this is the common basis of distinguishing them.

The solubility of hydroxides increases downward, and this is the reason why A_1 and B_1 formed white precipitates of magnesium hydroxide and calcium hydroxide respectively but C_1 does not because barium hydroxide is soluble in water.

Using ammonium hydroxide, which is a weak base, magnesium hydroxide precipitates out in test tube B_1 but the present hydroxide ions are not enough to precipitate out calcium hydroxide in B_2 which has a greater solubility than magnesium hydroxide.

In the end, sulphates become less soluble as we move down the group and on addition of aqueous sodium sulphate, C_3 remains a clear solution but B_3 becomes milky due to the formation of suspensions of $CaSO_4$ which is more soluble than $BaSO_4$ which precipitate out of C_3 as a white dense solid due to its highly limited solubility.

Evaluation

1. Explain the chemical test applied to identify the presence of magnesium and calcium ions by stating the reagent required and accompanying observations.
2. Given two unlabelled conical flasks containing potassium chloride and calcium chloride respectively. How can you proceed to label each container appropriately?

   Answer: *Take a sample from each flask, add sodium hydroxide. The sample from calcium chloride will give a white precipitate whereas the one from potassium chloride will remain colourless.*
3. You are provided with a solution of a mixture of magnesium nitrate and barium chloride. Suggest a simple experiment you can perform to remove magnesium ions from the mixture.

   Answer: *Treat the mixture with sodium hydroxide solution. Magnesium ions precipitate as magnesium hydroxide. Filtrate to remove the precipitate. The filtrate contains only barium ions.*

UNIT: 8 TRENDS OF CHEMICAL PROPERTIES OF GROUP 13 ELEMENTS AND THEIR COMPOUNDS

EXPERIMENT 8.1: REACTION OF ALUMINIUM WITH DIFFERENT ACIDS

Rationale

As a lightweight and strong metal, aluminium is used in power lines, construction of aircrafts, ships, trains, in high-rise buildings, window frames, household and industrial appliances, kitchen utensils due to the following property: durability, corrosive resistance, thermal and electrical conductivity among others. However, aluminium can be attacked by some concentrated acids.

Objective

Learners will be able to investigate the reaction of aluminium with concentrated hydrochloric and concentrated sulphuric acids

Required materials

Apparatus

- Bunsen burner
- Gas jar
- Thistle funnel
- Delivery tube
- Two stoppers with 2 holes
- Wooden splint
- Match box
- Two conical flasks

Chemicals

- Aluminium foil
- Concentrated HCl
- Concentrated H_2SO_4
- Solution of $KMnO_4$

Experimental set-up

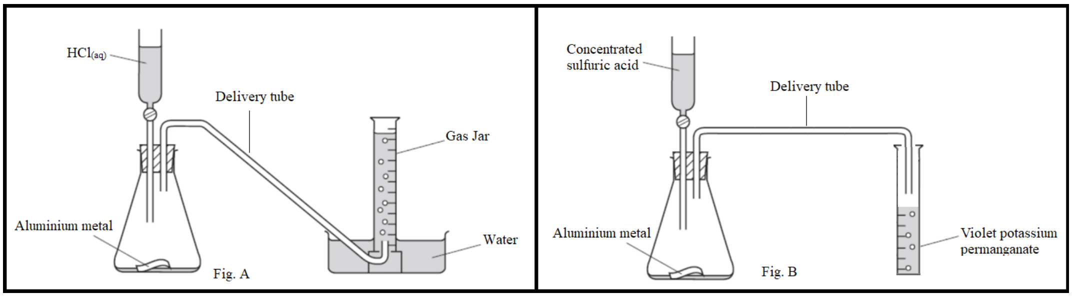


Figure 8.1: Reaction of Aluminium with acids (HCl, H_2SO_4)

Procedure

1. Put a small piece of aluminium foil into two different conical flasks and connect on each a thistle funnel and delivery tube as shown in figure A and B.
2. Pour moderately concentrated HCl in the thistle funnel as shown in Figure A. Open the tap of the thistle funnel and let hydrochloric acid flow on the aluminium metal in the flask.
3. Warm gently if the reaction is too slow. Record your observations.

   Answer: *A colourless gas is released.*
4. Collect the gas evolved and test it using a burning wooden splint.

   Answer: *The gas burns with a pop sound*
5. Pour concentrated sulphuric acid in the thistle funnel as shown in Figure B. Open the tap of the thistle funnel and let sulphuric acid flow on aluminium metal in the flask. Note your observation.

   Answer: *A colourless gas with a pungent odour is released*

6. Collect the gas evolved and test the gas using a violet solution of $KMnO_4$. Write down your observations.

   Answer: *The violet solution of $KMnO_4$ is decolorized by this gas.*

   Note: Aluminium is a corrosive resistant metal, but when it is attacked by some concentrated acids it forms salty solutions and releases some gases depending on the acids.

Interpretation of results and conclusion

Guiding questions

a) i)What happens when aluminium metal reacts with concentrated hydrochloric acid? Write the chemical equation for this reaction and name the formed products.

b) What happens when aluminium metal reacts with concentrated sulphuric acid? Write the chemical equation for this reaction and name the formed products.

Answer to guiding questions

When aluminium reacts with warm moderately concentrated hydrochloric acid, aluminium chloride and hydrogen gas are produced. The gas that burns with a pop sound is hydrogen

The chemical equation of reaction:

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

When aluminium reacts with concentrated sulphuric acid, aluminium sulphate, sulphur dioxide and water are formed. The gas which decolourises potassium dichromate is sulphur dioxide

The chemical equation of reaction:

$$2Al(s) + 6H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3SO_2(g) + 6H_2O(g)$$

Evaluation

1. By providing balanced chemical equations, differentiate and explain the reaction of HCl, H_2SO_4 and HNO_3 on aluminium metal.
2. Write a balanced chemical equation of the reaction between boron and hot concentrated sulphuric acid.

EXPERIMENT 8.2: REACTION OF ALUMINIUM WITH SODIUM HYDROXIDE

Rationale

As a lightweight and strong metal, aluminium is used in power lines, construction of aircrafts, ships, trains, in high-rise buildings, window frames, household and industrial appliances, kitchen utensils, due to the following property: durability, corrosive resistance, thermal and electrical conductivity among others. However, aluminium can be attacked by some concentrated alkaline solution.

Objective

Learners will be able to describe the reaction of aluminium with sodium hydroxide solution.

Required materials

Apparatus

- Borosilicate beaker
- Stirrer
- Thermometer
- Spatula

Chemicals

- Aluminium powder
- 40% sodium hydroxide solution
- Weighing balance

Procedure

1. Prepare 40% of sodium hydroxide by dissolving 40 g of sodium hydroxide in water to make 100 mL of the solution.
2. Take 0.5 g of aluminium powder into a borosilicate beaker.
3. Pour the solution of sodium hydroxide in the borosilicate beaker containing aluminium powder and allow the reaction to proceed for about 5 minutes.
4. Use a thermometer to record the temperature during the process. Record your observations.

   Answer: *Aluminium reacts vigorously with sodium hydroxide and the heat is released as the reaction occurs. A colourless gas has also evolved.*

Interpretation of results and conclusion

Guiding questions

a) What happens when aluminium is reacted with concentrated alkaline solution?

b) Write the chemical equation of chemical reaction between aluminium metal and concentrated sodium hydroxide and name the formed products.

c) What is the chemical test used to identify the gas released in question b?

Answer to guiding questions

When aluminium reacts with sodium hydroxide solution, sodium aluminate and hydrogen gas are formed with the heat production. This shows that the reaction is exothermic. The evolved gas burns with a pop sound. This gas is hydrogen.

The chemical equation of this reaction:

$$2Al(s) + 2NaOH(aq) + 6H_2O(l) \longrightarrow 2NaAl(OH)_4(aq) + 3H_2(g)$$

Evaluation

What are the products of the reaction between boron and concentrated potassium hydroxide? Write the chemical equation of the above reaction.

Answer: *The product s of the reaction are potassium borate and hydrogen gas*

The chemical equation of the reaction is $2B+6KOH \longrightarrow 2KBO_3+3H_2$

EXPERIMENT 8.3: REACTION OF ALUMINIUM OXIDE WITH ACIDS AND BASES

Rationale

Aluminium oxide is mainly used in making of refractories, ceramics, polishing and abrasive substances. It is also involved in the manufacture of zeolites used in water softeners. This experiment shows how aluminium oxide reacts with acids and alkalis.

Objective

Learners will be able to explain the amphoteric nature of aluminium oxide

Required materials

Apparatus

- Spatula
- Balance
- Droppers
- Beakers
- Test tubes
- Test tubes rack
- Bunsen burner
- Match box

Chemicals

- Aluminium oxide (Al_2O_3)
- Concentrated sodium hydroxide (3M)
- Dilute hydrochloric acid
- Aluminium hydroxide

Procedure

1. Take two test tubes and add 2 g of Al_2O_3 in each.
2. To the first test tube, add a few drops of hot concentrated sodium hydroxide solution.
3. Shake well. Record your observation?

   Answer: *Aluminium oxide reacts to give a colourless solution.*
4. To the second test tube, add dilute hydrochloric acid and heat. Note your observation.

   Answer: *A colourless solution is formed.*

Interpretation of results and conclusion

Guiding questions

1. Which products are formed when aluminium oxide reacts with:

   a) Hydrochloric acid

   b) Sodium hydroxide
2. Write a chemical equation of the above reactions.

Answer to guiding questions

Aluminium oxide reacts with hot concentrated sodium hydroxide solution to give sodium tetrahydroxo-aluminate.

$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \longrightarrow \underbrace{2NaAl(OH)_4(aq)}_{\text{Sodium tetrahydroxo-aluminate}}$$ Al_2O_3 acts as an acidicoxide

Aluminium oxide reacts with hot dilute hydrochloric acid to give aluminium chloride and water.

$$Al_2O_3(s) + 6HCl(aq) \longrightarrow \underbrace{2AlCl_3(aq)}_{\text{Aluminium trichloride}} + 3H_2O(l)$$ Al_2O_3 acts as a basic oxide

Aluminium oxide is amphoteric as it has both basic and acidic properties.

Evaluation

1. Why is aluminium oxide amphoteric?
2. Give two more examples of amphoteric substances and illustrate them using chemical equations.

EXPERIMENT 8.4: REACTION OF ALUMINIUM HYDROXIDE WITH ACIDS AND BASES

Rationale

Aluminium hydroxide is an amorphous white powder insoluble in water but soluble in alkaline and acidic solutions. It is used as an antacid to neutralize the acidity of the stomach when it has reached anomalous levels. This experiment explains how aluminium hydroxide behaves to neutralise hydrogen ions from acids and how it behaves in a basic medium.

Objective

Learners will be able to demonstrate the amphoteric nature of aluminium hydroxide

Required materials

Apparatus

- Spatula
- Balance
- Droppers
- Beakers
- Test tubes
- Bunsen burner

Chemicals

- Concentrated sodium hydroxide
- Dilute hydrochloric acid
- Aluminium hydroxide

Procedure

1. Take two clean test tubes and label them as A and B.
2. Add 2 g of $Al(OH)_3$ in each of the test tubes A and B.
3. Add a few drops of hot concentrated sodium hydroxide solution in test tube A and shake well. Write down your observations.

   Answer: *Aluminium hydroxide reacts with sodium hydroxide to give a colourless solution.*
4. Add dilute hydrochloric acid in test tube B and heat. Record your observation.

   Answer: *Aluminium hydroxide reacts with HCl(aq) to give a colourless solution as well.*

Interpretation of results and conclusion

Guiding question

What are the products of the reaction of aluminium hydroxide with

a) sodium hydroxide

b) hydrochloric acid?

Answer to guiding question

Aluminium hydroxide reacts with sodium hydroxide solution to give sodium tetrahydroxo-aluminate. The chemical equation of the reactions is

$$Al(OH)_3(aq) + NaOH(aq) \longrightarrow \underset{\text{Sodium tetrahydroxo-aluminate}}{Na[Al(OH)_4](aq)}$$ $Al(OH)_3$ reacts as an acidic oxide

Aluminium hydroxide reacts with dilute hydrochloric acid to give aluminium chloride solution. The chemical equation of the reaction is

$Al(OH)_3(s) + 3HCl(aq) \rightarrow Al\,Cl_3(aq) + 3H_2O(l)$ $Al(OH)_3$ reacts as a basic oxide

Both aluminium oxide and Aluminium hydroxide are amphoteric substances because they react with both strong bases and acids.

Evaluation

1. Aluminium hydroxide is used as an antacid in humans to reduce stomach acidity.
   a) Which acid is found in the human stomach?
   b) Write the chemical equation of reaction showing how aluminium hydroxide behaves to neutralise acidity in the stomach.
2. Aluminium acts as an amphoteric hydroxide. What does it mean? Explain your answer using an appropriate chemical reaction.
3. Write chemical equations of the reaction of beryllium hydroxide and
   a) Dilute sulphuric acid
   b) Sodium hydroxide

EXPERIMENT 8.5: CHEMICAL TEST FOR THE PRESENCE OF Al^{3+} ION IN THE SOLUTION

Rationale

Aluminium ion is a positively charged ion Al^{3+}. It is used in rechargeable aluminium -ion batteries. It is also used in hardness removal by using cation exchange resin in Al^{3+} form. This experiment determines the method of identifying the presence of aluminium ions in a given solution.

Objective

Learners will be able to test for the presence of Al^{3+} ions in a given solution.

Required materials

Apparatus

- Beaker
- Test tubes
- Droppers
- Test tube rack

Chemicals

- Aluminium salt solution
- Sodium hydroxide
- Ammonia solution

Procedure

1. Prepare two test tubes and label them as A and B
2. Pour about 2 mL of aluminium salt solution into each test tube (A and B)
3. Add dropwise sodium hydroxide solution until excess in test tube A. State your observation.

   Answer: *A white precipitate is formed and then dissolved, as sodium hydroxide is added in excess.*
4. Dropwise, add ammonia solution in a test tube B until excess. Record your observations.

   Answer: *A white precipitate is formed which is insoluble in excess ammonia solution.*

Interpretation of results and conclusion

Guiding question

How can you identify the presence of Al^{3+} ions in a given solution?

Answer to guiding question

When a solution of sodium hydroxide is added to the aqueous solution aluminium, a white precipitate of $Al(OH)_3$ is formed. $Al(OH)_3$ precipitate dissolves in excess of sodium hydroxide to give a soluble complex ion.

The ionic equation of this reaction:

$$Al^{3+}(aq) + 3Na(OH) \rightarrow \underset{\text{White ppt}}{Al(OH)_3\ (s)} + 3Na^{+}(aq)$$

$$Al(OH)_3(s) + OH\text{-}(aq) \rightarrow Al(OH)_4^-(aq)$$

With a solution of ammonia, aluminium ion Al^{3+} produces a white precipitate of $Al(OH)_3$, which is insoluble in excess ammonia solution.

The ionic equation of this reaction is

$$Al^{3+}(aq) + 3NH_4(OH) \rightarrow Al(OH)_3\ (s) + 3NH_4^{+}(aq)$$

$Al(OH)_3$ (s) + excess ammonia, no observable change.

Evaluation

1. A solution of water is suspected to contain aluminium chloride salt. How can you proceed to confirm the presence of aluminium ions in that solution? Illustrate your answer using appropriate chemical equations.
2. State the reagent that should be used to differentiate between a solution containing aluminium chloride salt and that containing calcium chloride salt. Illustrate your answer by chemical equations and expected observations

UNIT: 9

TRENDS IN CHEMICAL PROPERTIES OF GROUP 14 ELEMENTS AND THEIR COMPOUNDS

EXPERIMENT 9.1: REACTION OF CARBON WITH OXYGEN, CONCENTRATED ACIDS, DILUTE ACIDS AND HYDROXIDE.

Rationale

Carbon monoxide and carbon dioxide are produced when carbon is heated in the presence of oxygen, depending on the supply of oxygen. Carbon as a non-metallic element is the basic element of organic compounds and can be used as electrodes during electrolysis. Its compounds have many applications in daily life. Carbon dioxide is a raw material for photosynthesis. It is used in fire extinguishers and in soft drinks as conservative agents and as electrodes. Knowledge on carbon reactions helps to understand and differentiate products of carbon with oxygen, concentrated acids, dilute acids and hydroxide after performing experiments.

Objective

Learners will be able to demonstrate how carbon reacts with oxygen, dilute acids, concentrated acids and sodium hydroxide

Required materials

Apparatus

- Gas jar
- Boiling tube
- Test tube rack
- Bunsen burner
- Test tube holder
- deflagrating spoon

Chemicals

- Concentrated sulphuric acid
- Sodium hydroxide solution
- Charcoal
- Carbon powder

Procedure

1. Place a small piece of charcoal in 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: *Red-hot charcoal becomes brighter.*

2. Put 2 g of carbon powder into 3 different boiling tubes in the test tube rack.
3. Add 15 mL of concentrated sulphuric acid in the first boiling tube, 15 mL of dilute sulphuric acid in the second and 15 mL of sodium hydroxide in the third test tube and heat. What do you observe in all the three test tubes?

   Answer: *The colourless gas evolved in the first test tube. No observable change in the second and third test tubes.*

4. Test the gases produced by using lime water and potassium dichromate. Record your observation?

   Answer: *Lime water turns milky, and potassium dichromate turns green.*

Interpretation of results and conclusion

Guiding questions

1. What are the formed products when carbon reacts with oxygen?
2. Which gases are produced when carbon reacts with concentrated sulphuric acid?

Answer to guiding questions

When heated in the air, carbon reacts with oxygen to form carbon monoxide colourless gas (CO) in insufficient oxygen while there is formation of carbon dioxide (CO_2) in the presence of sufficient oxygen.

-Chemical equations of the reactions:

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

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

Carbon reacts with hot concentrated sulphuric acid to form carbon dioxide which turns lime water milky and sulphur dioxide having a pungent smell and turns orange potassium dichromate green.

-The chemical equation of the reaction is:

$$C(s) + H_2SO_4(l) \xrightarrow{\text{heat}} CO_2(g) + 2SO_2(g) + 2H_2O(l)$$

There is no reaction between carbon and dilute sulphuric acid and sodium hydroxide because carbon has a tendency to accept electrons and cannot displace hydrogen in dilute acid, while this is possible only in the concentration form.

Evaluation

1. What is the product formed when carbon reacts with concentrated sulphuric acid? Justify your answer.

   Answer: *When sulphuric acid reacts with carbon, it produces carbon dioxide, sulphur dioxide gas, and water.*

   The chemical equation of reaction:

   $C + H_2SO_4 \rightarrow CO_2 + H_2O + SO_2$

EXPERIMENT 9.2: REACTION OF TIN WITH OXYGEN, DILUTE ACIDS, CONCENTRATED ACID AND HYDROXIDES.

Rationale

Tin does not react with air under normal conditions. Tin is used mainly in tin plating, for making food cans, making soldiers, and bronze. It also acts as an electrode in batteries. Tin dissolves readily in concentrated acids and hot alkaline solutions, such as hot, concentrated potassium hydroxide. The reactivity of tin with oxygen, dilute acids, concentrated acid, and hydroxides will be investigated in this experiment.

Objective

Learners will be able to explain the reaction of tin with oxygen, dilute acid, concentrated acids, and hydroxide.

Required materials

Apparatus

- Gas jar
- Boiling tube
- Test tube rack
- Bunsen burner
- Test tube holder
- Deflagrating spoon

Chemicals

- Concentrated hydrochloric acid
- Sodium hydroxide solution
- Tin

Procedure

1. Place a small piece of tin in a deflagrating spoon and heat. Lower it in a gas jar of oxygen. What do you observe?

   Answer: *Tin burns when heated strongly to form a white solid.*

2. Put 2 g of tin into 3 different boiling tubes placed in a test tube rack.
3. Add 15 mL of concentrated hydrochloric acid in the first boiling tube, 15 mL of dilute hydrochloric acid in the second and 15 mL of sodium hydroxide in the third test tube. State your observation.

   Answer: *Tin reacts with concentrated hydrochloric acid to liberate a colourless gas which produces a pop sound when lit, but it does not readily react with dilute hydrochloric acid. Tin reacts with sodium hydroxide to form white precipitate.*

Interpretation of results and conclusion

Guiding questions

How does tin react with oxygen, dilute acids, concentrated acid and hydroxides? Write down the corresponding chemical equations.

Answer to guiding questions

Tin does not react with air under normal circumstances. Tin reacts with oxygen; when heated, it forms tin dioxide (SnO_2).
The chemical equation of reaction:

$Sn(s) + O_2(g) \rightarrow SnO_2(s)$

Note: Tin reacts with excess sodium hydroxide to form sodium stannite (Na_2SnO_2), sodium stannate (Na_2SnO_3) and hydrogen gas.

Chemical equation of reactions:

$Sn(s) + 2NaOH(aq) \rightarrow Na_2SnO_2(aq) + H_2(g)$

$Sn(s) + 2NaOH(aq) + H_2O(l) \rightarrow Na_2SnO_3(aq) + 2H_2(g)$

Typically, tin does not react with weak acids while strong acids can prompt chemical reactions. In this example, concentrated HCl acid combines with tin to form tin (II) chloride and hydrogen.

The chemical equation of reaction:

$Sn(s) + 2HCl \rightarrow SnCl_2(s) + H_2(g)$

With dilute acid there is no reaction, unless heat is applied to the solution.

Evaluation

1. What condition and product formed when tin reacts with oxygen?

   Answer: *Tin does not react with air under normal conditions. However, tin reacts with oxygen in the presence of heat to form tin oxide.*

   The equation of reaction is $Sn(s) + O_2(g) \rightarrow SnO_2(s)$

2. What is the product formed when tin reacts with hydroxide?

   Answer: *Tin forms a hydroxide compound called Tin (II) hydroxide:* $Sn(OH)_2$ *This molecule contains an octahedron of Sn atoms, each "capped" by oxide or hydroxide compounds.*

   The chemical equation of reaction:

   $Sn^{2+}(aq) + 2OH^-(aq) \rightarrow \underset{[white]}{Sn(OH)_2(s)}$

EXPERIMENT 9.3: REACTION OF LEAD WITH OXYGEN, DILUTE ACIDS, CONCENTRATED ACID AND HYDROXIDE

Rationale

The chemistry of lead contributes more to its reactivity. Lead is a reactive metal and reacts quickly with oxygen to form a layer of lead oxide. This prevents further reaction of lead with oxygen and other constituents found in air. Similar reaction is observed when lead metal is in contact with water to form a layer of lead carbonates, lead silicates, (depending on the dissolved ions in water) that protect the metal from further reactions. Because of this property, lead is mostly used for the lining of pipes that carry water or other liquids. It is also used in car batteries, making paints and as electrodes in electrolysis

Objective

Learners will be able to explain the reaction of lead with oxygen, dilute acids, concentrated acids and hydroxide

Required materials

Apparatus

- Gas jar
- Boiling tubes
- Test tube rack
- Bunsen burner
- Deflagrating spoon

Chemicals

- Concentrated hydrochloric acid
- Sodium hydroxide solution
- Lead

Procedure

1. Place a small piece of lead in a deflagrating spoon, heat it and then lower it in a gas jar of oxygen. Note your observation.

   Answer: *A reddish brown residue is formed when hot but later turns black and finally red.*

2. Put 2 g of lead into 3 different boiling tubes placed in a test tube rack.

3. Add 15 mL of concentrated hydrochloric acid in the first boiling, 15 mL of dilute hydrochloric acid in second and 15 mL of sodium hydroxide in the third test tube. Record your observations.

   Answer: *A colourless gas that produces a pop sound when lit in the first and third tube. No observable change in the second tube.*

Interpretation of results and conclusion

Guiding question

How does lead react with oxygen, dilute acids, concentrated acids, and hydroxide?

Answer to guiding question

Lead can form many different oxides. The most formed when it reacts with oxygen is lead (II) oxide, also known as lead monoxide. However, it may also form tri-lead tetroxide, sometimes called red lead, or lead (IV) oxide, also known as lead dioxide.

Lead + Oxygen → Lead (II) oxide

$2Pb + O_2 \rightarrow 2PbO$

Lead + Oxygen → Lead (IV) oxide

$Pb + O_2 \rightarrow PbO_2$

Lead + Oxygen → Tri-Lead tetroxide

$3Pb + 2O_2 \rightarrow Pb_3O_4$

Lead (Pb) reacts with concentrated hydrochloric acid (HCl) to form lead chloride ($PbCl_2$) and bubbles of hydrogen gas (H_2). The chemical equation can be written as:

$Pb(s) + 2HCl(aq) \rightarrow PbCl_2\,(aq) + H_2(g)$

Lead is above hydrogen in the activity series, but it does not react with dilute hydrochloric acid. This is because it forms an insoluble coating of lead chloride. Therefore, further reaction is prevented with the formation of the coating layer of lead chloride.

Evaluation

1. Why does not lead react with dilute acid?

   Answer: *This is because lead forms an insoluble coating of lead chloride which prevent further reactions*

2. Lead forms different oxides with oxygen. Write 2 different balanced chemical equations between lead and oxygen.

   Answer: *See interpretation*

EXPERIMENT 9.4: REACTION OF GROUP 14 OXIDES WITH WATER, ACIDS, AND BASES

Rationale

This experiment introduces learners to the behaviour of some group 14 oxides when treated with water, acids, and bases. The oxides of group 14 elements have many applications in our daily life. Lead oxide is used in storage batteries, pigments, ceramics, and glass industry. Tin oxide finds its potential application in catalysis, electrocatalysis, solar energy conversion etc. Silicon oxide is commonly used in the production of elemental silicon which is used as a fining agent in juice, beer, and wine. It is also used in pharmaceuticals for making tablets.

Objective

Learners will be able to explain the reaction of group 14 oxides with water, acids, and bases.

Required materials

Apparatus

- Test tubes
- Measuring cylinder
- Electronic balance
- Spatula

Chemicals

- Concentrated hydrochloric acid
- Sodium hydroxide solution
- Water
- Silicon oxide
- Lead oxide
- Tin oxide

Procedure

1. Prepare three dry test tubes and put about 1 g of lead oxide into each test tube
2. Add 2 mL concentrated hydrochloric acid solution in the first test tube and shake. What do you observe?

   Answer: *Lead oxide dissolves in concentrated hydrochloric acid solution*
3. Add 2 mL sodium hydroxide solution in the second test tube and shake. What do you observe?

   Answer: *Lead oxide dissolves in sodium hydroxide solution*
4. Add 2 mL of water in the third test tube and shake. What do you observe?

   Answer: *Lead oxide does not dissolve in water*
5. Repeat steps 1 to 4 by using silicon dioxide and Tin oxide instead of lead oxide.

   Answer: *Silicon dioxide and tin oxide do not dissolve in water but they dissolve in acid and base*

Interpretation of results and conclusion

Guiding questions

1. How do lead, silicon and tin oxides react with water, acids, and bases? Explain.
2. Explain the amphoteric nature of tin (IV) oxide by using chemical reactions.
3. Explain the amphoteric nature of lead (II) oxide by using chemical reactions.

Answer to guiding questions

Reaction of SiO_2 with water:

Silicon dioxide does not react with water.

Reaction of SiO_2 with acids:

Silicon dioxide reacts with HF acid only:

$SiO_2(s) + 4HF(aq) \rightarrow SiF_4(aq) + 2H_2O(l)$

This reaction is the one used to write on the glass, for example on the windscreen of vehicles.

Reaction of SiO_2 with bases: Silicon dioxide reacts with hot concentrated NaOH solution.

$SiO_2(s) + 2OH^-(aq) \rightarrow SiO_3^{2-}(aq) + H_2O(l)$

Reaction of SnO_2 with water: SnO_2 does not react with water.

Reaction of SnO and SnO_2 with acids:

SnO_2 also reacts with hot concentrated HCl acid.

$SnO_2(s) + 4HCl(conc.) \xrightarrow{heat} SnCl_4(aq) + 2H_2O(l)$

SnO also reacts with concentrated HCl acid.

$SnO(s) + 2HCl(aq) \rightarrow SnCl_2(aq) + H_2O(l)$

Reaction of SnO_2 and SnO with bases

SnO_2 reacts with concentrated NaOH to produce a stannate (IV) ion:

$SnO_2(s) + 2OH^-\ conc \xrightarrow{heat} SnCl_4(aq) + 2H_2O(l)$

SnO reacts with dilute NaOH to produce a tin sodium stannate

$SnO + 2NaOH \rightarrow Na_2SnO_2 + H_2O$

Lead (II) oxide is an amphoteric oxide. It behaves as a base when it reacts with acids and it behaves as an acid, when it reacts with bases.

$PbO(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2O(l)$

$PbO(s) + 2OH^-(aq) + H_2O(l) \rightarrow Pb(OH)_4(aq)$

$PbO(s) + H_2O(l)$ No reaction

Evaluation

1. Write the chemical equations to show the amphoteric nature of lead (II) oxide.

   Answer: *See interpretation*

2. Discuss the chemical reactions of tin (II) oxide and lead (II) oxide with sulphuric acid.

   Answer: *See interpretation*

3. The windscreens of the vehicle are made up of silicon dioxide. Explain how one could write words on them without using the painting process.

   Answer: *See interpretation*

4. Write the chemical equation of the reaction between lead (IV) oxide and:

   a) Concentrated hydrochloric acid

   b) Concentrated sodium hydroxide

EXPERIMENT 9.5: REACTION OF GROUP 14 CHLORIDES WITH WATER, ACIDS, AND BASES.

Rationale

The reaction of lead (II) chloride with sulphuric acid produces lead (II) sulphate that is used in lead–acid batteries, as paint pigments, and as laboratory reagent. In this experiment, we shall investigate the behaviour of lead (II) chloride in the presence of water, acidic and alkaline solutions.

Objective

Learners will be able to demonstrate the chemical reaction of group 14 chlorides with water, acids, and bases.

Required materials

Apparatus

- Test tubes
- Test tube racks
- Spatula

Chemicals

- Concentrated sulphuric acid solution
- Sodium hydroxide solution
- Lead (II) chloride
- Silicon (II)chloride
- Tin (II) chloride
- Water

Procedure

1. Prepare four dry test tubes and put in about 1 g of lead (II) chloride in each.
2. In the first test tube, add 2 mL of concentrated sulphuric acid solution and shake.
3. In the second test tube, add 2 mL of sodium hydroxide solution and shake.
4. In the third test tube, add 2 mL of water and shake.
5. In the fourth test tube, add 2 mL of hot water and shake.
6. Repeat by using silicon and tin chlorides. What do you observe?

   Answer: *Group 14 chlorides dissolve in sulphuric acid, sodium hydroxide and hot water but they do not dissolve in cold water.*

Interpretation of results and conclusion

Guiding questions

1. How do lead (II) chloride), silicon(IV) chloride), tin(IV) chloride) react with

   a) sulphuric acid

   b) sodium hydroxide

   c) water

Write down chemical equations of the above reactions.

Answer to guiding questions

Lead (II) chloride reacts with sulphuric acid to produce lead (II) sulphate and hydrogen chloride. $PbCl_2 + H_2SO_4 \rightarrow PbSO_4 + 2HCl$

Lead (II) chloride does not react with cold water, but it reacts with hot water to produce lead (II) chloride-hydroxide and hydrogen chloride.

$PbCl_2 + H_2O \rightarrow Pb\,(OH)Cl + HCl$

The tetrachloride of group 14 elements are colourless liquids apart from $PbCl_4$ which is a yellow liquid $PbCl_4$ and $SnCl_4$ are unstable, and they decompose to $PbCl_2$ and Cl_2, $SnCl_2$ and Cl_2 respectively.

Tin (IV) chloride reacts with concentrated sulphuric acid to produce tin (IV) sulphate and hydrogen chloride. The reaction takes place in a boiling solution.

$SnCl_4 + 2H_2SO_4 \rightarrow Sn(SO_4)_2 + 4HCl$

Tin (IV) chloride reacts with sodium hydroxide to produce tin (IV) hydroxide and sodium chloride.

$SnCl_4 + 4NaOH \rightarrow Sn(OH)_4 + 4NaCl$

Silicon (IV) chloride hydrolyses in water producing silicon dioxide and white fumes of hydrogen chloride gas:

$SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + HCl(g)$

$SiCl_4(l) + 2H_2O(l) \rightarrow SO_2(s) + HCl(g)$

Evaluation

1. Explain why lead (IV) chloride is unstable.
2. Write chemical equation of the reaction between silicon chloride with
   a) water
   b) sodium hydrochloric acid

EXPERIMENT 9.6: CHEMICAL TEST OF TIN IONS

Rationale

The identification of ions is an important activity as it helps to know qualitatively the composition of a given mixture. By knowing its composition, one can handle it safely and efficiently. In this experiment we have to deal with the test for the presence of tin (II) ion in solution.

Objective

Learners will be able to demonstrate how tin ions can be identified in solution.

Required materials

Apparatus

- Test tubes
- Dropper
- Test tubes rack

Chemicals

- Tin (II) chloride
- Sodium hydroxide solution
- Ammonia solution
- Measuring cylinder
- Mercury nitrate

Procedure

1. Take two different test tubes and pour 2 mL of tin chloride in each.
2. Add a few drops of sodium hydroxide solution in the first test tube. Write down your observations.

   Answer: A white precipitate is formed.
3. Add again a few drops of sodium hydroxide solution. What do you observe?

   Answer: *By adding a few drops, the white precipitate dissolves.*

4. Add a few drops of ammonia solution in the second test tube and record your observation.

   Answer: *a white precipitate is formed.*

5. Add again a few drops of ammonia. Write down your observations?

   Answer: *by adding a few drops, the white precipitate does not dissolve in excess.*

6. Add three drops of mercury nitrate In the third test tube, and state your observation.

   Answer: *white precipitate which turns grey bluish on standing in the air*.

Interpretation of results and conclusion

Guiding questions

1. How do tin ions react with sodium hydroxide in excess. Write the chemical equation of the reaction.
2. How can you confirm the presence of tin (II) ions in a given solution?

Answer to guiding questions

Sodium hydroxide and ammonia solutions are reagents used to detect the possible presence of tin ions in solution. The reaction of tin (II) chloride and sodium hydroxide produces a white precipitate of hydrated tin (II) oxide which is soluble in excess. The chemical equations of the reactions that take place are:

$$SnCl_2(aq) + 2NaOH\ (aq) \rightarrow SnO{\cdot}H_2O\ (s) + 2NaCl\ (aq)$$

$$SnO{\cdot}H_2O\ (s) + NaOH(aq) \rightarrow NaSn(OH)_3(aq)$$

The reaction of tin (II) chloride and ammonia solution gives a white precipitate which is insoluble in excess.

The presence of Sn^{2+} ion is confirmed by the addition of a colourless solution of mercury nitrate to a solution suspected to contain Sn^{2+}ion. The positive test gives a *white precipitate which turns grey bluish on standing in the air*.

Evaluation

1. Describe a test that may be used to identify the presence of tin (II) ion.
2. Write a chemical equation of the reaction of silicon oxide with sodium hydroxide.

EXPERIMENT 9.7: CHEMICAL TEST OF CARBONATES (CO_3^{2-})

Rationale

Carbonates find their application in many processes like in glasses making, in paper industries, in water treatment, soap and detergent making, in chalks productions, and in medical plasters. Carbonates contain the carbonate ion (CO_3^{2-}). They are tested by the action of dilute acid on the carbonate compounds. The reaction releases carbon dioxide gas which forms a white precipitate of calcium carbonate when bubbled through limewater. This test is characteristic for carbonates that distinguish them from other salts.

Objective

Learners will be able to perform chemical test of carbonates (CO_3^{2-})

Required materials

Apparatus

- Test tubes
- Dropper
- Delivery tube

Chemicals

- Sodium carbonate salt solution
- Dilute hydrochloric acid
- Calcium hydroxide

Experimental set-up

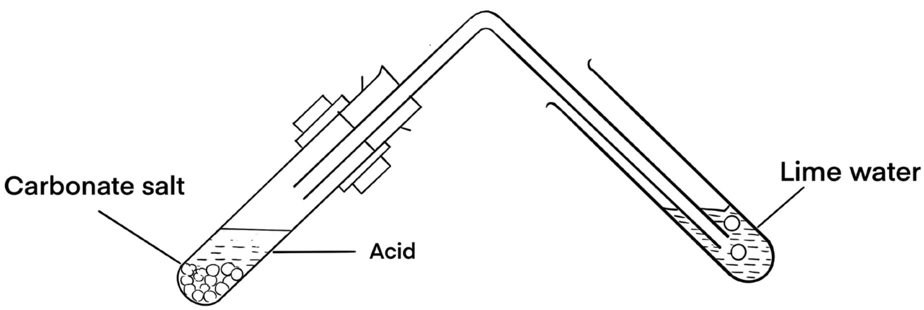


Figure 9.1: Test for carbonates

Procedure

1. Pour into the test tube a few drops of sodium carbonate solution.
2. Add a few drops of dilute hydrochloric acid and by using a delivery tube, connect it to another test tube containing lime water. Record your observations.

   Answer: *The gas evolved turns lime water milky*

Interpretation of results and conclusion

Guiding question

What chemical test could be used to confirm the presence of carbonate ions in solution?

Answer to guiding question:

The confirmation test of the presence of carbonate ions in solution is done by reacting the solution to be tested by dilute acid. If a carbonate compound is present, the effervescence forms. The gas **e**volved is carbon dioxide which turns lime water milky.

Evaluation

2 mL of dilute nitric acid was added to 0.5 g of zinc carbonate in a test tube. The produced gas was then dissolved in calcium hydroxide solution.

a) State the observations made.

b) Write stoichiometric chemical equations for the reactions that occur.

c) Draw a setup of the apparatus that can be used to perform this experiment.

UNIT: 10

TRENDS IN CHEMICAL PROPERTIES OF GROUP 15 ELEMENTS AND THEIR COMPOUNDS

EXPERIMENT 10.1: REACTION OF PHOSPHORUS WITH OXYGEN

Rationale

Phosphorus has many allotropes. One of them, red phosphorus, is highly toxic and bursts into flames when exposed to air. Red phosphorus is used in the striking surfaces of matches because the phosphorus ignites with friction. The study of the reaction of phosphorus with oxygen will help to be aware when dealing with these allotropes. This experiment illustrates the reaction of red phosphorus with oxygen.

Objective

Learners will be able to explain the reaction of red phosphorus with oxygen.

Required materials

Apparatus

- Gas jar
- Bunsen burner
- Match box
- Deflagrating spoon
- Test tubes
- Test tubes rack

Chemicals

- Red phosphorus
- Water

Procedure

1. Place a small piece of red phosphorus on a deflagrating spoon.
2. Heat it then lower it in a gas jar of air. Record your observation

   Answer: *A white solid is formed.*
3. Dissolve a small amount of the product in water contained in a test tube.
4. Test the resulting solution using blue and red litmus papers. state your observations.

   Answer: *Blue litmus paper turns red and red litmus paper remains unchanged.*

Interpretation of results and conclusion

Guiding questions

1. How does red phosphorus react with oxygen? Write a balanced chemical equation of the reaction.
2. What is the formed product when compound obtained in 1. reacts with water? Write the balanced corresponding chemical equation.

Answer to guiding questions:

Hot red phosphorus burns in air and two phosphorus oxides can be produced.

In limited oxygen, phosphorus (III) oxide is formed

$P_4(s) + 3O_2(g) \rightarrow 2P_2O_3(s)$

In sufficient oxygen it forms phosphorus pentoxide.

$P_4(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$

When phosphorus pentoxide reacts with water, an acidic solution which turns a blue litmus paper red is obtained.

The chemical equation of the reaction:

$$P_2O_5(s) + 3H_2O(l) \rightarrow 2H_3PO_4(aq)$$

All group 15 elements form two types of oxides: E_2O_3 and E_2O_5. The oxide in the higher oxidation state of the element is more acidic than that of lower oxidation state. Their acidic character decreases down the group. Group 15 oxides of phosphorus are acidic.

Evaluation

1. Write a balanced chemical equation of the reaction of red phosphorus with oxygen and suggest how the products can be tested.
2. Write balanced chemical equations of the reaction of oxygen with: a) nitrogen b) Arsenic c) antimony.

EXPERIMENT 10.2: REACTION OF NITRIC ACID WITH METALS AND NON-METALS

Rationale

Nitric acid is a highly corrosive mineral acid commonly used in the production of ammonium nitrate which is a constituent of most fertilizers. The acid is also used in the production of explosives as well as in the oxidation of metals. Being a powerful oxidizing acid, nitric acid reacts with many organic materials, and the reactions may be explosive. This experiment provides useful information on its reactivity with metals and non- metals. Therefore, safety precautions need to be recognised while working with this acid.

Objective

Learners will be able to demonstrate and explain the reactions of nitric acid with metals and non-metals.

Required materials

Apparatus

- Spatula
- Test tubes
- Bunsen burner
- Test tube holder

Chemicals

- Concentrated nitric acid(70%)
- Copper turnings
- Magnesium ribbon
- Zinc powder
- Carbon (or charcoal)
- Sulphur powder
- Red and blue litmus paper
- Bromine

Caution

- Work in a well-ventilated room.
- Avoid big yield of the gas when HNO_3 reacts with non-metals and metals. The gas is toxic. Therefore, use a small amount of the reagents.
- Concentrated acidic is corrosive; much care must be taken in using it.
- The reactions of nitric acid with non- metals can be violent, and nitrogen dioxide is very poisonous.

Procedure

1. Pour 5 mL of concentrated nitric acid in a clean test tube.
2. Add a small piece of magnesium to the test tube. What do you observe?

   Answer: *brown fumes of a pungent smell are released.*
3. Test the gas produced with wet blue and red litmus papers. What do you observe?

   Answer: *Blue litmus paper turns red and the colour of red litmus paper does not change.*
4. Repeat the experiment using copper turnings.

   Answer: *Reddish-brown fumes are observed.*
5. Repeat the experiment using zinc powder. What do you observe?

   Answer: *Reddish-brown fumes are observed.*
6. Put one spatula of carbon powder into a beaker; add about 5 mL of concentrated nitric acid. If no reaction is taking place, gently heat the mixture.
7. Test the gas produced with a wet blue and red litmus paper and then a wet paper which has been dipped in lime water. What do you observe?

   Answer: *-A reddish-brown gas has evolved. Blue litmus paper turns red. Red litmus paper does not change.*

   -On the surface of the paper that has been dipped in lime water a white solid is formed.

8. Repeat the procedure in 5 and 6 using sulphur powder, add two or three drops of bromine to catalyse the reaction. State your observations.

Answer: *Reddish-brown fumes are observed*.

Interpretation of results and conclusion

Guiding question

How does nitric acid react with metals and non-metals?

Answer to guiding question

Reaction of nitric acid with metals

* **Reaction with Magnesium**

Concentrated nitric acid reacts with magnesium oxidizing it to magnesium (II) nitrate, $Mg(NO_3)_2$, nitrogen dioxide and water.

$$Mg(s) + HNO_3(l) \rightarrow Mg(NO_3)_2(aq) + NO_2(g) + H_2O(l)$$

* **Reaction with copper**

Concentrated nitric acid reacts with copper to give copper (II) nitrate, $Cu(NO_3)_2$, nitrogen (IV) oxide and water.

$$Cu(s) + 4HNO_3(l) \rightarrow \underset{\text{Blue solution}}{Cu(NO_3)_2(aq)} + \underset{\text{brown gas}}{2NO_2(g)} + 2H_2O(l)$$

* **Reaction with Zinc**

Concentrated nitric acid reacts with zinc to give zinc nitrate, nitrogen dioxide and water.

$$Zn(s) + HNO_3(l) \rightarrow Zn(NO_3)_2(aq) + NO_2(g) + H_2O(l)$$

The reddish-brown gas which turns blue litmus paper red is nitrogen dioxide.

Note: Unlike other acids, nitric acid rarely gives off hydrogen gas with metals except for very dilute nitric acid (about 1%) which reacts with magnesium by liberating hydrogen gas.

$$Mg(s) + 2HNO_3(dil.) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$$

Reactions of nitric acid with non-metals

Fuming nitric acid gives the most vigorous reaction giving off brown fumes of nitrogen dioxide.

* **Reaction with carbon**

Hot concentrated nitric acid reacts with carbon to give off carbon dioxide, nitrogen dioxide and water.

$C(s) + 4HNO_3 (l) \rightarrow CO_2 (g) + 4NO_2 (g) + 2H_2O(l)$

The gas that reacts with lime water to give a white solid is carbon dioxide.

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

* **Reaction with Sulphur**

Concentrated or fuming nitric acid reacts with sulphur to give off reddish-brown fumes of nitrogen (IV) oxide (NO_2). Two or three drops of bromine can be added to catalyse the reaction.

$S(s) + 6HNO_3(l) \rightarrow H_2SO_4 (aq) + 6NO_2 (g) + 2H_2O(l)$

Evaluation

1. Write chemical equations for reactions of zinc, copper, and magnesium with concentrated and dilute nitric acid.

   Answer: *see interpretation*

2. Give chemical equations for the reactions of concentrated nitric acids with non-metals (carbon and sulphur).

   Answer: *see interpretation*

EXPERIMENT 10.3: REACTION OF PHOSPHORIC ACID (H_3PO_4) WITH METALS AND BASES

Rationale

Phosphoric acid is a triprotic acid that can be obtained by reaction of tetraphosphorus decaoxide (P_4O_{10}) with water. In industry, it is prepared by reaction between phosphates and concentrated sulphuric acid. As an acid, it reacts with some metals and alkaline solutions to give phosphates some of which are used as fertilisers. We consider in this experiment the reactions of phosphoric acid with magnesium, lead, copper, and calcium hydroxide.

Objective

Learners will be able to explain the reaction of phosphoric acid (H_3PO_4) with metals and bases

Required materials

Apparatus

- Test tubes
- Measuring cylinder
- Test tubes rack

Chemicals

- Concentrated phosphoric acid
- Magnesium
- Copper
- Lead
- Calcium hydroxide

Procedure

1. Pour 5 mL of concentrated phosphoric acid in a clean test tube.
2. Add a small piece of magnesium to the test tube. What do you observe?

   ***Answer**: Release of a colourless gas.*
3. Test the gas produced using a lit splint. What are your observations?

   ***Answer**: The gas released burns with a pop sound.*
4. Repeat the experiment using copper turnings. What are your observations?

   ***Answer**: No noticeable observation.*
5. Repeat the experiment using lead metal. What do you observe?

   ***Answer**: Release of a colourless gas.*
6. Take another test tube containing 5 mL of concentrated phosphoric acid and add in a small amount of calcium hydroxide. What do you observe?

   Answer: White precipitate is formed.

Interpretation of results and conclusion

Guiding questions

1. What is the identity of the products formed when phosphoric acid reacts with lead, copper, and magnesium?
2. What is the type of reaction when calcium hydroxide reacts with phosphoric acid?

Answer to guiding questions

I. Reaction with metals

Phosphoric acid reacts with magnesium to form their corresponding metal phosphates and hydrogen.

$3Mg(s) + 2H_3PO_4 \rightarrow Mg_3(PO_4)_2(s) + 3H_2(g)$

$3Pb_{(s)} + 2H_3PO_4 \rightarrow Pb_3(PO_4)_{2(s)} + 3H_{2(g)}$

The hydrogen burns with a pop sound.

Copper does not react with phosphoric acid under normal conditions. Copper is under hydrogen in the reactivity series. Therefore, it cannot displace hydrogen from acids. In addition, phosphoric acid is a weak oxidising agent, and it cannot oxidise copper.

II. Reaction with bases

Phosphoric acid reacts with calcium hydroxide to form salt and water. This is an acid-base neutralisation reaction accompanied with evolution of heat.

$$H_3PO_4(l) + 3Ca(OH)_2(s) \rightarrow Ca_3(PO_4)_2(aq) + 3H_2O(l)$$

Evaluation

1. Predict the product of the reaction between phosphoric acid and aluminium metal. Write the corresponding chemical equation.
2. Phosphoric acid reacts with sodium hydroxide producing a salt. Give the name and three uses of the salt which is formed.

EXPERIMENT 10.4: REACTION OF PHOSPHORUS OXIDES/ CHLORIDES WITH WATER

Rationale

The reaction of phosphorus oxides or phosphorus chlorides produce acids (phosphoric acid and phosphorous acid) which turn blue litmus paper red. Phosphorus pentoxide is used as a strong drying and dehydrating agent. Phosphorus trichloride is used in gasoline additives, textiles finishing, and to make pesticides, catalysts, and plasticisers. This experiment highlights the chemical behaviour of these oxides/chlorides with water.

Objective

Learners will be able to demonstrate the reaction of phosphorus oxides/ chlorides with water.

Required materials

Apparatus

- Beaker
- Spatula

Chemicals

- Phosphorus oxides
- Phosphorus chloride
- Distilled water
- Litmus papers

Procedure

Take one full spatula of P_2O_3, P_2O_5, PCl_3, PCl_5 and put them in 4 different beakers respectively and then add water.

Test the solution using blue and red litmus papers. Write down your observation.

Answer: *The blue litmus paper turns red and there is no change with red litmus paper.*

Interpretation of results and conclusion

Guiding question

How do phosphorus oxides/chlorides react with water?

Answer to guiding question:

When phosphorus oxides/chlorides react with water, an acid which turns blue litmus paper red is produced..

Chemical reaction of P_2O_3 with water:

$P_2O_3(s) + 3H_2O(l) \rightarrow 2H_3PO_3(aq)$

Reaction of P_2O_5 with water:

$P_2O_5(s) + 3H_2O(l) \rightarrow 2H_3PO_4(aq)$

Reaction of PCl_3 with water

$PCl_3(l) + 3H_2O \rightarrow H_3PO_3(aq) + 3HCl(aq)$

Reaction of PCl_5 with water:

$PCl_5(l) + 4H_2O\ (l) \rightarrow H_3PO_3(aq) + 5HCl(aq)$

Evaluation

1. Predict the formed products when $P_2O_{3,}$ P_2O_5, PCl_3, PCl_5 react with water. Write their corresponding balanced chemical equations.

   Answer: *See interpretation*

2. Between PCl_3 and $SbCl_3$, which one is more reactive with water? Explain your answer

   Answer: *Down the group, the reactivity decreases due to the increase in stability. Therefore $PCl_{3,}$ undergoes hydrolysis reaction more than $SbCl_{3,}$ hence PCl_3 reacts more with water than PCl_3 which undergoes dissolution.*

EXPERIMENT 10.5: REACTION OF PHOSPHATES WITH SULPHURIC ACID

Rationale

Phosphorous provided to the plants in the form of phosphates is essential in plants for root formation, seed development and plant maturation. Calcium phosphate makes the main structure of bones and teeth. Phosphates can also be used to produce phosphoric acid used in the food industry, cosmetics, and animal feed. In this experiment sulphuric acid is used to study the behaviour of phosphates in acidic media.

Objective

Learners will be able to conduct a reaction of phosphates with sulphuric acid.

Required materials

Apparatus

- Spatula
- Pyrex/borosilicate beaker
- Dropper
- Conical flask
- Funnel
- Filter papers
- Thermometer

Chemicals

- Blue litmus paper
- Concentrated sulphuric acid
- Calcium phosphate

Procedure

1. Take a spatula full of calcium phosphate and put it in Pyrex/ borosilicate beaker
2. Add few drops of concentrated sulphuric acid
3. Heat the content of the beaker up to 180^{0}C (measure the temperature with the help of a thermometer). Record your observations.

Answer: *A solid substance is observed at the bottom of the beaker.*

4. Filter the content of the beaker to remove the insoluble substance.
5. Test the filtrate using blue litmus paper. What do you observe?

   Answer: *Blue litmus red.*

Interpretation of results and conclusion

Guiding questions

What are the reaction products between calcium phosphate and sulphuric acid? Write a balanced chemical equation.

Answer to guiding questions

Heating a mixture of calcium phosphate and concentrated sulphuric acid at a temperature slightly less than 180 ºC produces phosphoric acid and the insoluble calcium sulphate which is filtered off.

A blue litmus paper dipped in the resulting solution turns red which confirms the presence of an acid in the solution. Hence, this is a simple process that can be used to prepare phosphoric acid.

$Ca_3(PO_4)_2(s) + 3H_2SO_4(l) \rightarrow 3CaSO_4(s) + 2H_3PO_4(aq)$

Evaluation

Write a chemical equation of reaction required to synthesize phosphoric acid.

EXPERIMENT 10.6: IDENTIFICATION TEST OF PHOSPHATE AND NITRATE IONS

Rationale

Nitrates and phosphate are naturally occurring compounds on earth that are composed of nitrogen, phosphorus, and oxygen elements. Testing nitrates and phosphates helps to identify their presence in foods, soil, and water. Phosphate tests are very useful for measuring phosphate levels in people who are malnourished. This experiment deals with the test for the presence of phosphate and nitrate ions in solution.

Objective

Learners will be able to identify phosphate and nitrate ions in a solution.

Required materials

Apparatus

- Test tubes
- Bunsen burner
- Spatula
- Test tube holder
- Dropper

Chemicals

- Copper
- Sodium nitrate (NaNO3)
- Concentrated sulphuric acid
- Concentrated nitric acid,
- Suspected salt of phosphate
- Ammonium molybdate solution

Procedure

1. Put a full spatula of $NaNO_3$ into a test tube.
2. Add some drops of concentrated sulphuric acid then heat for about 30 seconds. State your observations.

   Answer: *Brown gas on the walls of the test tube is observed.*

3. Add a piece of copper metal in a test tube. What do you observe?

   Answer: *Intensity of brown gas increases.*

4. Put a full spatula of a phosphate in another clean test tube.
5. Add some drops of concentrated nitric acid using a dropper.
6. Boil the contents of the test tube over the Bunsen burner.
7. Add excess of ammonium molybdate solution using another dropper. Record your observations.

   Answer: *A bright yellow precipitate is observed.*

Interpretation of results and conclusion

Guiding questions

1. How is nitrate ion identified?
2. What is the confirmatory test of phosphate ions?

Answer to guiding questions

Sulphuric acid reacts with the nitrate ion to form nitric acid. Nitric acid (very strong oxidizing agent) then reacts with the copper turnings to form nitric oxide. Nitric oxide is thus oxidized to nitrogen dioxide which is brown gas as it is shown by this reaction

$2NaNO_3(s) + H_2SO_4(l) \rightarrow Na_2SO_4(aq) + 2HNO_3(aq)$

$Cu + 4\ HNO_3 \rightarrow Cu(NO_3)_2 + 2\ NO_2 + 2H_2O$

The presence of phosphate ion, PO_4^{3-} is indicated by a bright yellow precipitate of ammonium molybdophosphate, $(NH_4)_3P(Mo_3O_{10})_4$, formed upon treatment of the test solution with ammonium molybdate–nitric acid reagent.

$PO_4^{3-}(aq) + 12MoO_4^{2-}(aq) + 3NH_4^+(aq) + 24H^+(aq) \rightarrow (NH_4)_3P(Mo_3O_{10})_4(s) + 12H_2O(l)$

Evaluation

1. Explain how nitrates ions can be tested using sulphuric acid.

   Answer: *Sulphuric acid can displace HNO_3 from any nitrate, then HNO_3 formed decomposes slightly on heat and oxidises copper to release NO_2 gas.*

2. Explain how phosphates ions can be identified.

Answer: *A small portion of the sample is acidified with concentrated nitric acid and ammonium molybdate is added. The formation of a bright yellow precipitate layer of ammonium phosphomolybdate indicates the presence of phosphate ions. Gentle heating can hasten the appearance of the precipitate.*

UNIT: 11

TRENDS OF CHEMICAL PROPERTIES OF GROUP 16 ELEMENTS AND THEIR COMPOUNDS

EXPERIMENT 11.1: REACTION OF CONCENTRATED SULPHURIC ACID WITH POTASSIUM HALIDES

Rationale

The sulphuric acid is used in the preparation of hydrogen halides because it is a non-volatile acid with a high boiling point (356°C). This acid reacts with any of metal halide salts to form hydrogen halides by displacement reaction. Sulphuric acid is used in fertiliser making, insecticides, pharmaceutical drugs, paints, pigments, and as electrolyte in car batteries. The experiment demonstrates its reaction with potassium halides.

Objective

Learners will be able to describe the reaction of concentrated sulphuric acid with potassium chloride, potassium bromide and potassium iodide.

Required materials

Apparatus

- Spatula
- Test tubes
- Test tube rack
- Test tube holder

Chemicals

- Concentrated sulphuric acid
- Potassium iodide
- Potassium bromide
- Potassium chloride

Caution

Work in a well-ventilated room

Procedure

1. Take three clean test tubes and label them as A, B and C.
2. Put a half spatula of potassium chloride in test tube A, potassium bromide test tube B and potassium iodide in test tube C.
3. Add 4 drops of concentrated sulphuric acid in each test tube. Record your observations?

Answer:

Solid halide	*Observations*	*Further observations*
KCl	*White fumes*	No further reaction
KBr	White fumes	Brown fumes and choking smell
KI	White fumes	Purple fumes, choking smell and a rotten egg smell

Interpretation of results and conclusion

Guiding questions

1. What are the products of the reactions of sulphuric acid with:
   a) Potassium chloride?
   b) Potassium bromide?
   c) Potassium iodide?

Answer of guiding questions

Concentrated sulphuric acid reacts with potassium chloride to produce white fumes of HCl.

$H_2SO_4(l) + KCl(s) \rightarrow HCl(g) + KHSO_4(s)$

The reaction of potassium bromide with concentrated sulphuric acid gives white fumes of HBr. After a whilethe concentrated sulphuric acid oxidises and decomposes HBr into a reddish-brown gas of bromine while sulphuric acid is itself reduced to sulphur dioxide, choking gas.

$H_2SO_4(l) + KBr(s) \rightarrow HBr(g) + KHSO_4(s)$

$2HBr(g) + H_2SO_4(l) \rightarrow Br_2(g) + SO_2(g) + 2H_2O(l)$

Potassium iodide reacts with concentrated sulphuric acid to produce white fumes of HI.

$H_2SO_4(l) + KI(s) \rightarrow HI(g) + KHSO_4(s)$

Sulphuric acid oxidises the hydrogen iodide to purple elemental iodine, and it is itself reduced to sulphur dioxide (a choking gas), deposits of solid sulphur and hydrogen sulphide (with a smell of rotten eggs).

The chemical equations of the reactions taking place

$2HI(g) + H_2SO_4(l) \rightarrow I_2(g) + SO_2(g) + 2H_2O(l)$

$6HI(g) + H_2SO_4(l) \rightarrow 3I_2(g) + S(s) + 4H_2O(l)$

$8HI(g) + H_2SO_4(l) \rightarrow 4I_2(g) + H_2S(g) + 4H_2O(l)$

Evaluation

1. Explain the reason why sulphuric acid is used in preparation of halogen halides.

   Answer: *The Sulphuric acid is used in the preparation of hydrogen halides because it is a non-volatile acid*

2. There are some traces of iodine during the reaction of iodide salts and the concentrated sulphuric acid. Why are no chlorine traces observed during this such reaction?

   Answer: *The reducing power of the halide ions increases down the group which is the reason why the chloride ion cannot reduce sulphur in the* SO_4^{2-}

EXPERIMENT 11.2: REACTION OF SULPHITE IONS WITH DILUTE ACIDS

Rationale

Any sulphite reacts with dilute acids to produce a soluble salt and a gas is released while sulphates don't react with acids, that is why we are going to use sodium sulphite and hydrochloric acid to produce a gas. Sulphites are used as food additives to prevent microbial spoilage, preserve food colour and to maintain the potency of certain medications. This experiment deals with the reaction of sodium sulphite with dilute acids.

Objective

Learners will be able to explain the reaction of sulphite ions with dilute acids.

Required materials

Apparatus

- Spatula
- Round bottomed flask
- Washing bottle
- Gas jar
- Delivery tubes
- Bunsen burner
- Match box
- Retort stand and accessories
- 2 stoppers with 2 holes
- Thistle funnel

Chemicals

- Dilute sulphuric acid, H_2SO_4
- Sodium sulphite, Na_2SO_3
- Dilute hydrochloric acid
- Potassium dichromate

Experimental set-up

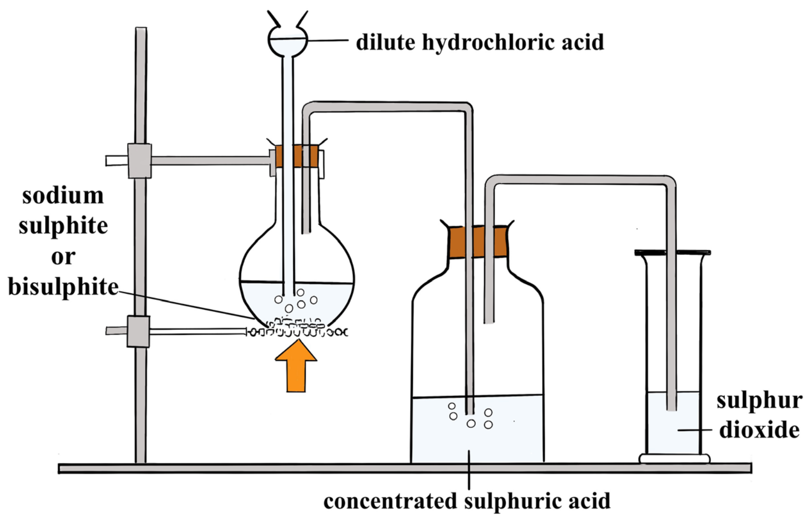


Figure 11.1: Reaction of Sodium sulphite with dilute HCl

Procedure

1. Put a half spatula of sodium sulphite in a round bottomed flask.
2. Set the apparatus as shown in figure 11.1
3. Pour 10 mL of acidified potassium dichromate in a gas jar.
4. Add five drops of dilute hydrochloric acid in a round bottomed flask containing sodium sulphite and then heat. Write down your observations.

   ***Answer**: A colourless gas is evolved which turns yellow acidified potassium dichromate solution green .*

Interpretation of results and conclusion

Guiding questions

1. What are the products of the reaction between sodium sulphite and hydrochloric acid ?
2. How is the evolved gas tested?
3. What is the role of concentrated sulphuric acid in this experiment?

Answer to guiding questions

Dilute hydrochloric acid reacts with sodium sulphite to produce SO_2 gas.

The chemical equation for the reaction:

$Na_2SO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g)$

Concentrated sulphuric acid is used to dry obtained sulphur dioxide gas. Sulphur dioxide colourless gas is tested using yellow acidified potassium dichromate solution that turns green.

Evaluation

1. a) Write a balanced chemical equation between potassium sulphite and dilute hydrochloric acid?

   b) How the gas produced in (a), if any, can be tested?

EXPERIMENT 11.3: ACTION OF HEAT ON SULPHATES

Rationale

Some salts decompose easily on heat, but sulphate salts resist heat. However, they immediately lose their water of crystallisation. Sulphate products are used in the production of fertilizers, chemicals, dyes, glass, paper, soaps, textiles, and insecticides. This experiment demonstrates the effect of heat on sulphates.

Objective

Learners will be able to describe the action of heat on sulphates.

Required materials

Apparatus

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

Chemicals

- Hydrated copper (II) sulphate
- Hydrated iron (II) sulphate

Experimental set-up

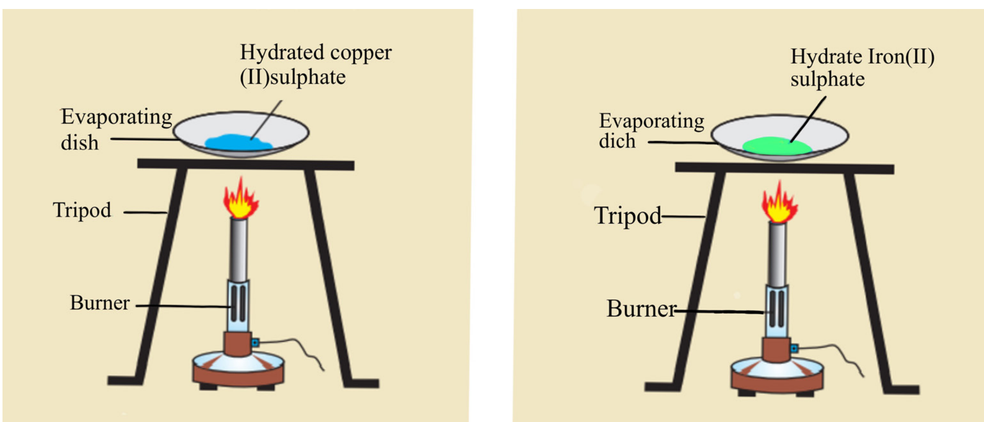


Figure 11. 3: Action of heat on sulphates

Caution

Put on safety glasses, do not look directly into the evaporating dish, keep distance from the burner and do not touch the hot dish.

Procedure

1. Set up all apparatus as shown in the figure 11.3.
2. Put 2 g of copper (II) sulphate crystals on an evaporating dish and heat gently. What do you observe?

   Answer: *Blue copper sulphate crystals become white.*
3. Continue heating strongly for some minutes. Note your observations.

   Answer: *No change is observed on further heating.*
4. Repeat the experiment using iron (II) sulphate crystals.

   Answer: *Green crystals of iron (II) sulphate change into white solid and no observable change takes place on further strong heating.*

Interpretation of results and conclusion

Guiding question

What do you observe when hydrated copper sulphates and iron (II) sulphates are heated?

Answer to guiding question

On heating:

i. iron (II) sulphate, first loses its water of crystallization and the original green crystals are converted into a white anhydrous powder which turns brown when exposed to air.

ii. hydrated copper (II) sulphate also loses its water of crystallization and original blue crystals change into a white anhydrous copper sulphate powder.

Evaluation

Describe the colour change of hydrated copper sulphate and Iron (II)sulphate before and after gentle heat.

Answer: *Before heating, Copper sulphate is blue. On gentle heating, it loses its water of crystallisation and becomes white powder.*

Iron II sulphate is green solid. On heating, iron II sulphate becomes white cream powder.

EXPERIMENT 11.4: REACTION OF IODINE WITH THIOSULPHATE IONS ($S_2O_3^{2-}$)

Rationale

Thiosulphate ions combined with sodium are widely used in analysing mineral composition, gold mining, and chlorination of water. It is used in medicine for the treatment of cyanide poisoning. This experiment demonstrates one of the reactions that take place in analytical chemistry, iodometry, which involves thiosulphate ions and iodine. Iodine is an important chemical in our body of which its deficiency causes the swelling up of thyroid gland (known as goitre).

Objective

Learners will be able to explain the reaction between iodine and thiosulphate ions.

Required materials

Apparatus

- Test tube
- Dropper
- Test tubes rack

Chemicals

- Iodine solution
- Sodium thiosulphate solution (0.1M)

Procedure

1. Prepare iodine solution by dissolving 10g KI and 5 g I_2 in 100mL of distilled water.
2. Add 2 mL of iodine solution in a test tube.
3. Add dropwise a sodium thiosulphate solution. Note your oberservations.

   Answer: *Brown colour of iodine disappears after a while.*

Interpretation of results and conclusion

Guiding questions

1. Why does the iodine solution become colourless when sodium thiosulphate solution is added to it?
2. Does iodine react with sodium thiosulphate? If yes, when write a balanced chemical equation for the reaction

Answer to guiding questions

Thiosulphate ions react with molecular iodine through an oxidation-reduction reaction to produce iodide ions and tetrathionate ions in solution. Both ions produced are colourless whereas the initial solution is brown due to the presence of dissolved iodine. The chemical equation for the reaction is

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

Thiosulphate ions are reducing agents while the iodine molecule is an oxidizing agent.

Evaluation

1. The reaction between iodine and thiosulphate ions is an oxidation-reduction reaction. Explain.
2. What do you observe when starch is added to an iodine solution?
3. Why is iodine mixed with potassium iodide (KI) before its reaction with thiosulphate ions?

   Answer: *To make ease the dissolution of iodine in water*

EXPERIMENT 11.5: IDENTIFICATION TESTS FOR SULPHITE AND SULPHATE IONS

Rationale

Sulphates are used in cement production, plaster formation, and detergents production. Sulphites are naturally found in some foods and act as food preservatives. Though they are both very important in our everyday life, it needs more attention while handling them. Otherwise, it can lead to serious chemical accidents and health hazards. In this experiment, identification of sulphite and sulphate ions will be performed.

Objective

Learners will be able to identify sulphate and sulphite ions in a given solution.

Required materials

Apparatus

- Test tubes
- Droppers
- Test tube rack

Chemicals

- Sodium sulphite
- Distilled water
- Sodium sulphate
- Barium nitrate
- Dilute nitric acid

Procedure

1. Put two test tubes in a test tube rack and label them as A and B.
2. Add 2 mL of sodium sulphate solution in test tube A and 2 mL of sodium sulphite solution in test tube B.
3. Add four drops of barium nitrate solution to each of the test tubes. What do you observe?

   Answer: *A white precipitate is formed in both test tubes.*
4. Add 4 drops of dilute nitric acid to the resultant mixture in each test tube and shake. Note your observations?

   Answer*: In test tube A there is no observable change whereas in test tube B the precipitate dissolves.*

Interpretation of results and conclusion

Guiding question

How can sulphate be differentiated from sulphite ions?Justify your answer by using balanced chemical equations.

Answer to guiding question

Both sulphates and sulphites form white precipitates in the presence of $Ba^{2+}(aq)$.

$SO_3^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_3(s)$ white precipitate

$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$ white precipitate

In the presence of dilute nitric acid, $BaSO_3$ dissolves while $BaSO_4$ does not dissolve.

$BaSO_3(aq) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + SO_2(g) + H_2O(l)$

The sulphate ion is a weak base which does not react with acids whereas the sulphite ion reacts with them because it is a strong base.

Generally, the sulphite ion reacts as follows:

$SO_3^{2-} + 2H^+ \rightarrow SO_2 + H_2O$

Evaluation

1. Why is it important to test for the presence of sulphite and sulphate ion?
2. Explain why barium sulphite dissolves in nitric acid whereas barium sulphate does not.

UNIT: 12

TRENDS OF CHEMICAL PROPERTIES OF GROUP 17 ELEMENTS AND THEIR COMPOUNDS

EXPERIMENT 12.1: PREPARATION AND CHEMICAL TEST OF CHLORINE

Rationale

Chlorine is an important chemical used in water treatment, bleaching substances manufacturing, and insecticide production. However, it is very toxic and should be handled with care. Due to its high reactivity in nature, it is always found in combined forms. In this experiment a simple method of producing and testing chlorine in the laboratory is described.

Objective

Learners will be able to describe how to prepare and test chlorine gas in the laboratory.

Required materials

Apparatus

- Round bottomed flask
- Conical flask
- Bunsen burner
- 2 delivery tubes
- Gas jar
- Thistle funnel

Chemicals

- Concentrated sulphuric acid
- Concentrated hydrochloric acid
- Potassium permanganate
- Blue and red litmus papers

Experimental set-up

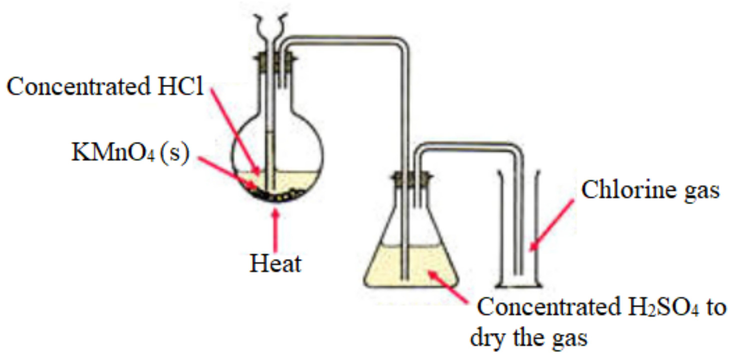


Figure 12.1: Preparation of chlorine gas

Caution

This should be done by the teacher only in a fume cupboard.

Procedure

1. Set up the pieces of apparatus as shown in the figure 12.1.
2. Put some crystals of potassium permanganate in the round-bottomed flask and connect a thistle funnel containing concentrated HCl.
3. Allow the HCl from the thistle funnel to flow into the round-bottomed flask and heat the mixture. What do you observe?

   Answer: *A yellowish green gas is produced in the gas jar.*
4. Put some concentrated H_2SO_4 in the conical flask to dry the gas.
5. Test the gas with wet blue and red litmus papers. What do you observe?

   Answer: *The red litmus paper does not change but the blue one turns red, and it bleaches out after a while.*

Interpretation of results and conclusion

Guiding question

How is chlorine prepared and tested in the laboratory?Support your answers with chemical equations.

Answer to guiding questions:

Hydrochloric acid reacts with potassium permanganate to produce various products including chlorine gas. This is because $KMnO_4$ is a strong oxidizing agent which oxidizes chloride ions. The chemical equation of the reaction:

$$2KMnO_4(s) + 16HCl(aq) \rightarrow 2KCl(aq) + 2MnCl_2(aq) + 5Cl_2(g) + 8H_2O(l)$$

Chlorine gas is proved to be an acidic gas which also has bleaching properties. However, it does not bleach dry litmus papers because the bleaching ion, OCl^-, is only formed when chlorine dissolves in water as shown by the next chemical equation.

$$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$$

Evaluation

1. Chlorine is prepared by heating a mixture of manganese dioxide and hydrochloric acid.
   a) Explain why the above mixture must be heated whereas when potassium manganate (VII) is used there is no need of heat.
   b) Write a balanced chemical equation of the above reactions.

EXPERIMENT 12.2: PREPARATION OF BROMINE AND IODINE

Rationale

Bromine is a toxic red-brown oily liquid with a sharp smell. It is used in many areas such as agricultural chemicals, dyestuffs, insecticides, pharmaceuticals, fire retardants and chemical intermediates. However, some uses are being abandoned due to environmental issues.

Iodine, a black shiny crystalline solid sublimes to a purple vapour when gently heated. It has many uses in photographic chemicals, disinfectants, pharmaceuticals, dyes, and printing inks. It is also added to table salt in the form of iodide to prevent risks of getting goitre, the swelling up of thyroid gland caused by iodine deficiency.

As halogens are very reactive and always combined with other elements in nature.In this experiment, simple methods of preparing them in the laboratory are elaborated.

Objective

Learners will be able to describe how bromine and iodine are prepared in the laboratory.

Required materials

Apparatus

- Test tubes

Chemicals

- Sodium/Potassium bromide
- Sodium/Potassium iodine
- Concentrated H_2SO_4

Procedure

1. Take two test tubes and label them as A and B.
2. Put some potassium bromide crystals in test tube A and potassium iodide in test tube B.
3. Add some drops of concentrated sulphuric acid to each test tube. Record your observations.

   Answer: *In the test tube A white fumes that turn reddish-brown after a while are produced. In test tube B white fumes which turn purple gas which solidifies on the wall of the tube is formed.*

Interpretation of results and conclusion

Guiding question

How are bromine and iodine prepared in the laboratory? Support your answers with chemical equations.

Answer to guiding question

The reaction of potassium bromide and concentrated sulphuric acid gives white fumes of HBr and after a while, the concentrated sulphuric acid oxidizes HBr into a reddish-brown bromine gas.

The chemical equations of the reactions taking place are:

$H_2SO_4(l) + KBr(s) \rightarrow HBr\ (g) + KHSO_4(s)$

$2HBr\ (g) + H_2SO_4\ (l) \rightarrow Br_2\ (g) + SO_2\ (g) + 2H_2O\ (l)$

Sodium/potassium iodide reacts with concentrated sulphuric acid to produce white fumes of HI.

Immediately excess sulphuric acid oxidizes the hydrogen iodide to purple iodine.

The chemical equations of the reactions taking place are:

$H_2SO_4(l) + KI(s) \rightarrow HI\ (g) + KHSO_4(s)$

$8HI\ (g) + H_2SO_4\ (l) \rightarrow 4I_2\ (g) + H_2S\ (g) + 4H_2O\ (l)$

Sulphuric acid is an oxidizing agent and halide ions are reducing agents. The reducing power of the halide ions increases down the group which is the reason why the chloride ion cannot be oxidized to produce chlorine whereas the bromide and the iodide ions can be easily reduced to elemental bromine and iodine respectively.

Evaluation

1. Bromine and iodine are prepared by reacting a metal bromide or iodide with concentrated sulphuric acid. Why does this method not apply for the preparation of chlorine?
2. What precautions must be taken during the preparation of these halogens (bromine and iodine)?

   Answer: *Work under a fume hood.*

EXPERIMENT 12.3: DISPLACEMENT OF THE IODIDE AND BROMIDE IONS BY CHLORINE

Rationale

We have already seen that the halogens are very reactive that they are never found free in nature. However, their reactivity changes as we move down in the group, and this can be used to extract elemental less reactive halogens from solutions of their soluble salts using more reactive ones. In this experiment bromine and iodine are produced using chlorine.

Objective

Learners will be able to explain the displacement of the iodide and bromide ions by chlorine.

Required materials

Apparatus

- Test tube
- Bunsen burner
- Spatula
- Beaker
- 2 Retort stands
- Delivery tube
- Stopper with one hole
- Round bottomed flask

Chemicals

- Sodium chloride crystals
- Concentrated H_2SO_4
- Sodium bromide
- Potassium iodide
- Potassium permanganate
- Starch
- Water
- Litmus paper

Experimental set-up

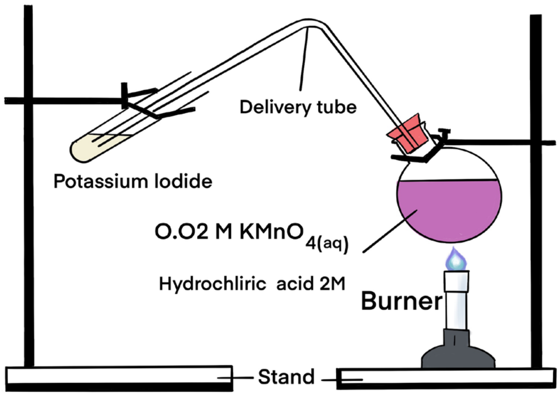


Figure 12.1: Displacement of the iodide and bromide ions by chlorine

Procedure

1. Dissolve 2 g of potassium iodide in a beaker containing 120 mL of water.
2. Pour 20 mL of a 0.02M potassium permanganate solution into a round bottomed flask.
3. Pour 20 mL of concentrated hydrochloric acid (2M HCl) into potassium permanganate in the round bottomed flask.
4. Add 10 mL of potassium iodide solution in the boiling tube.
5. Insert the delivery tube into a boiling tube containing potassium iodide solution as shown in the set-up diagram (Figure 12.1).
6. Heat the mixture in the round-bottomed flask until there is a colour change in the test tube. What do you observe?

   Answer: *The colour changes from colourless to brown.*
7. Add a few drops of starch indicator to the resultant solution in the boiling tube. What do you observe?

   Answer: *The brown colour of the solution changes to dark blue.*
8. Repeat steps 1 to 6 using potassium bromide. What do you observe?

   Answer: *In the test tube the colour changes from colourless to reddish brown.*

Interpretation of results and conclusion

Guiding questions

1. How bromine and iodine are displaced from their salts by chlorine?. Briefly explain.
2. Write balanced chemical equations of the above reactions.

Answer to guiding questions

Chlorine displaces iodine from potassium iodide solution. Iodine slightly dissolves in water to give a dark brown solution, and turns the starch indicator dark blue.

The greyish-black residue is due to the formation of solid iodine solid.

$2KI(aq) + Cl_2(g) \rightarrow 2KCl\,(aq) + I_2\,(aq)$

Colourless Brown colour

$I_2\,(aq) + Starch \rightarrow$ Starch-Iodine complex

Brown Dark blue

Starch-Iodine complex (on standing) $\rightarrow I_2(s)$ + Starch

dark blue deep violet

Chlorine also displaces bromide ions from their solution to produce aqueous bromine.

$2KBr\,(aq) + Cl_2\,(g) \rightarrow 2KCl\,(aq) + Br_2\,(aq)$

In the test tube the colour changes from colourless to reddish brown as bromine is formed.

Note: The reducing power of halide ions increases down the group whereas the oxidizing power decreases as we move down. This is the reason why each halide can be oxidized by the one which is above it.

Evaluation

1. Simply explain how bromine and iodine are displaced by chlorine.
2. Complete each of the following chemical equations;
   a) $NaBr(aq) + Cl_2(g) \rightarrow$
   b) $NaBr(aq) + I_2(aq) \rightarrow$

EXPERIMENT 12.4: TEST FOR THE PRESENCE OF HALIDE IONS IN A SOLUTION

Rationale

Many inorganic compounds are white solids regardless of the different ions making them. Due to this, it is always imperative to know the exact composition of a compound to avoid possible misuse of chemicals and accidents and/or health hazards that may be caused. This experiment explains how to distinguish chloride, bromide, and iodide ions.

Objective

Learners will be able to test the presence of halide ions from solutions of their salts.

Required materials

Apparatus

- Test tubes
- Test tube holders
- Droppers
- Measuring cylinder

Chemicals

- Silver nitrate
- Distilled water
- Potassium chloride
- Potassium bromide
- Potassium iodide

Procedure

1. Avail three well cleaned test tubes and label them as A, B, and C.
2. Add about 3 mL of a solution containing potassium chloride in a test tube A, 3 mL of a solution containing potassium bromide in a test tube B, 3 mL of a solution containing potassium iodide in a test tube C.
3. Add 5 drops of 0.1M silver nitrate solution in each test tube. What do you observe?

   Answer: *In test tube A, a white precipitate is formed, in test tube B a white cream is formed, in test tube C a pale yellow precipitate is formed.*

4. Then in each test tube, add a few drops of ammonia solution and then in excess. What do you observe?

Answer: In test tube A a white precipitate dissolves to form a colourless solution. In test tube B, persist. In test tube C, pale yellow persists.

Interpretation of results and conclusion

Guiding questions

1. How can you distinguish chloride, bromide and iodide ions in a given solution? support your answer by chemical equations.

2. **Answer to guiding questions:**

Halide ion	*Observation on addition of silver nitrate*	*Observation on addition of excess ammonia solution*	*Equation for reaction*
Cl^-	*White precipitate*	*White precipitate soluble*	$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ white precipitate $AgCl(s) + 2NH_3(aq) \rightarrow [Ag(NH_3)]^+(aq)$ Colourless solution
Br^-	*Pale-yellow precipitate*	*Pale-yellow precipitate insoluble*	$Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$ cream precipitate
I^-	*Yellow precipitate*	*Yellow precipitate insoluble*	$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$ pale-yellow precipitate

Evaluation

1. Describe what is observed when chloride, bromide and iodide are tested using a soluble salt of lead (II) and write the appropriate chemical equations.
2. For testing the presence of halide ions using silver nitrate solution, it is advised to use distilled water instead of tap water. Suggest reasons.

UNIT: 15

FACTORS THAT AFFECT THE CHEMICAL EQUILIBRIUM

EXPERIMENT 15.1: ILLUSTRATION OF THE REVERSIBILITY OF SOME CHEMICAL REACTIONS

Rationale

Chemical equilibrium is very common in many industrial processes such as the Haber process for the production of ammonia and contact process for the production of sulphuric acid. These processes involve reactions that take place in both directions.

Many chemical reactions do not go up to completion because after the formation of some products, the latter start to undergo a reverse reaction and make reactants. This experiment is designed to give concrete examples of a reversible reaction.

Objective

Learners will be able to demonstrate the reversibility of some chemical reactions.

Required materials

Apparatus

- Droppers
- Test tubes
- Dropper

Chemicals

- 10% NaOH solution
- 0.1M K_2CrO_4
- 6M HNO_3 solution

Procedure

1. Place 3 mL of 0.1 M potassium chromate in a test tube.
2. Add an equal amount of 0.6 M nitric acid solution. What do you observe?

   Answer: a yellow solution turns into orange.
3. Then add 10% sodium hydroxide dropwise. What do you observe?

   Answer: *Orange solution turns into yellow.*

Interpretation of results and conclusion

Guiding question

What does reversibility of chemical reactions mean? Explain with supporting examples.

Answer to guiding questions

When aqueous nitric acid is added in a test tube containing potassium chromate solution, the colour changes from yellow to orange.

$2Cr_2O_4^{2-}(aq) + 2H^+(aq) \rightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$

Yellow Orange

On the other hand, the addition of NaOH(aq) in the orange solution removes hydrogen ions and dichromate ions start to change back into chromate ions which are yellow.

$H^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$

The existence of both hydrogen ions and hydroxide ions in the reacting mixture makes the process reversible. Eventually both processes will end up taking place at the same speed with no clear observable change, but still changes will be taking place.

Evaluation

1. Write the balanced chemical equations for nitric acid and potassium chromate.

   Answer: $2K_2CrO_4 + 4HNO_3 \rightarrow 4KNO_3 + H_2Cr_2O_7 + H_2O$

2. Explain why the colour of potassium chromate changed from orange to yellow after the addition of sodium hydroxide.

   Answer: *The addition of sodium hydroxide in the orange solution removes hydrogen ions and dichromate ions, coloured orange, start to change back into chromate ions which are yellow.*

UNIT: 17 REDUCTION AND OXIDATION REACTIONS

EXPERIMENT 17.1: ILLUSTRATION OF OXIDATION REDUCTION REACTIONS

Rationale

Oxidation and reduction reactions are very common, and they take place in processes like rusting, respiration, extraction of metals among others. This experiment shows how more reactive metals can displace less reactive ones from solutions of their salts through oxidation and reduction reactions.

Objective

Learners will be able to illustrate oxidation reduction reactions.

Required materials

Apparatus

– Beaker

Chemicals

– Zinc

– Copper nitrate solution

Procedure

1. Clean a metallic strip of zinc properly.
2. Place a strip of metallic zinc in an aqueous solution of copper nitrate for about one hour. What do you observe?

   Answer: *Red-brown particles form at the bottom of the beaker as the blue colour of copper sulphate fade*

Interpretation of results and conclusion

Guiding question

Is the reaction between zinc metal and copper nitrate oxidation or a reduction reaction? Explain.

Answer to guiding questions

You may notice that the size of the zinc strip reduces, and solid red-brown particles are deposited at the bottom of the beaker. Moreover, the blue colour of the copper (II) sulphate solution fades.The reaction between metallic zinc and the aqueous solution of copper nitrate is

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

In this reaction, zinc has lost electrons to form Zn^{2+} and, therefore, zinc is oxidised.

Evidently, now if zinc is oxidised by releasing electrons. Copper ions are reduced by gaining electrons from the zinc.

Evaluation

1. Why does the blue colour of copper nitrate solution fade out slowly when it is reacting with zinc metal?

   Answer*: See interpretation*

2. Write the equation of reaction between metallic zinc and the aqueous solution of copper nitrate.

   Answer*: See interpretation*

3. Explain why zinc is a reducing agent while copper (II) ion is an oxidizing agent.

   Answer: *See interpretation*

UNIT: 18

ENERGY CHANGES AND ENERGY PROFILE DIAGRAMS FOR CHEMICAL REACTIONS

EXPERIMENT 18.1: EXPERIMENT TO VERIFY ENERGY CHANGES DURING A CHEMICAL REACTION

Rationale

Chemical reactions are always accompanied by either a release or absorption of heat. Heat releasing reactions/processes are very useful in our daily activities such as cooking, car driving, and home warming. Our body temperature is kept constant by heat releasing reactions that take place in our body. On the other hand, heat absorbing reactions/processes are used in food preservation by refrigeration. This experiment gives a clear example of heat releasing and absorption reactions.

Objective

Learners will be able to demonstrate how energy changes during a chemical reaction.

Required materials

Apparatus

- Thermometer
- Plastic beaker
- Spatula
- Stirring rod

Chemicals

- Vinegar
- Baking soda
- Calcium chloride
- Water

Procedure

1. Avail two beakers, and label them A and B.
2. Add 10 mL of water to a beaker A and place a thermometer in the water. Record the initial temperature (T_i).
3. Add spatula end full of calcium chloride to the water in beaker A and stir carefully using a stirring rod. Record the final temperature (T_f).

   Answer: *The temperature increases.*
4. Place the thermometer in the beaker B containing vinegar. Record the initial temperature.
5. Add an endful spatula of baking soda and gently stir using a stirring rod. Record the final temperature. What do you observe?

   Answer: *The temperature decreases.*

Interpretation of results and conclusion

Guiding questions

1. Why does the temperature change when $CaCl_2$ is added to water?
2. Why does the temperature change when baking soda is added to vinegar?
3. What do you call a chemical reaction accompanied by an increase in temperature?

Answer to guiding questions

Calcium chloride is a chemical compound made up of calcium ions and chloride ions. The ions are held together by an ionic bond. Mixing calcium chloride with water is an exothermic reaction, which means that the combination of the two substances releases heat and the temperature increases.

It took energy to break the baking soda and vinegar apart and energy is absorbed when the carbon dioxide, sodium acetate, and water are formed. Since more energy is needed to break the baking soda and vinegar apart, the temperature reduces. This reaction is called an endothermic reaction.

Evaluation

1. Provide two more examples of exothermic reactions.

   Answer: *snow formation in clouds, burning of a candle, rusting of iron, burning of sugar, formation of ion pairs, reaction of strong acid and water, water and calcium chloride, etc.*

2. Identify the differences between endothermic and exothermic reactions in terms of energy change during chemical reactions.

References

1. Electrolytes and Nonelectrolytes: Types, Differences & Examples | AESL (aakash.ac.in)
2. https://www.aakash.ac.in/important-concepts/chemistry/electrolytes-and-non-electrolytes