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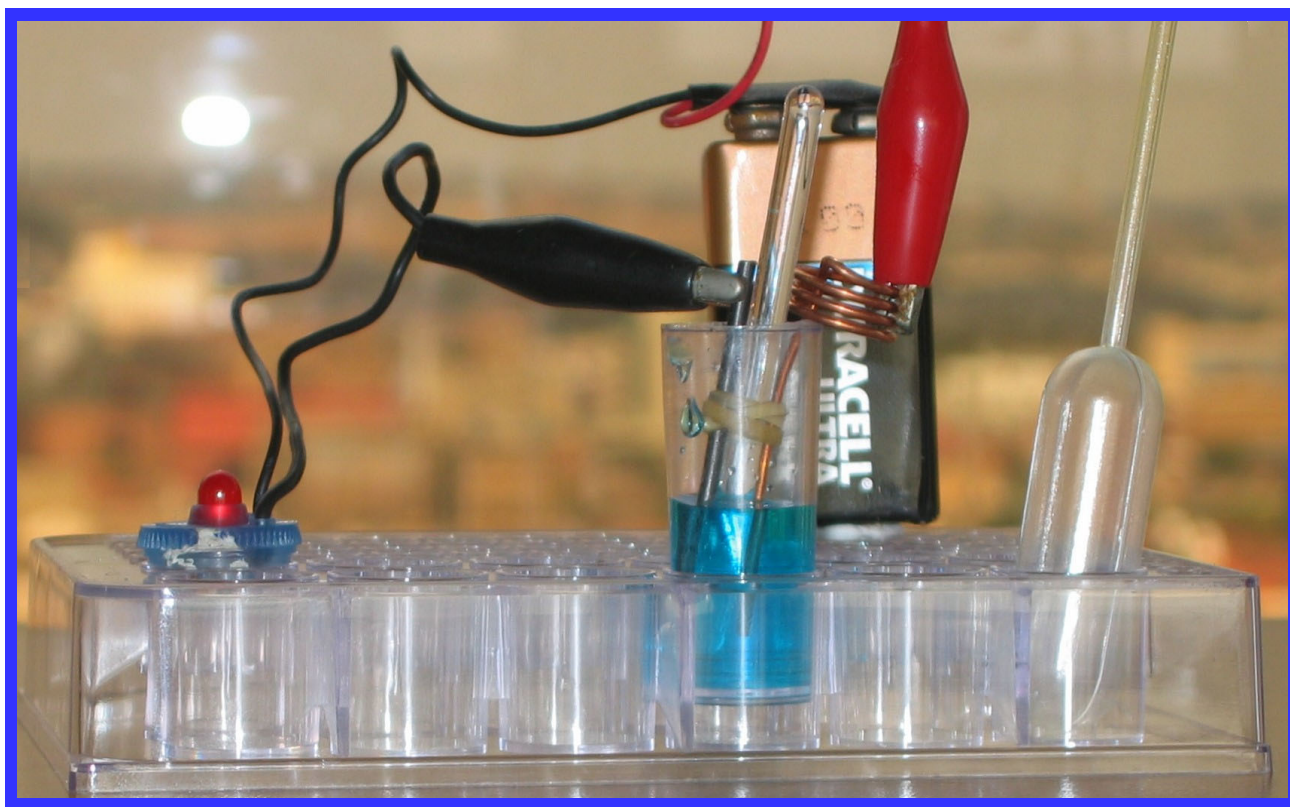
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ADVANCED TEACHING PACKAGES

MICROELECTROCHEMISTRY EXPERIMENTS

Manual for Teachers - First Edition



Compiled by Beverly Bell, Bina Akoobhai
Edited by Prof. JD Bradley
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United Nations Educational,
Scientific and Cultural Organization



The UNESCO-Associated Centre
for Microscience Experiments,
RADMASTE Centre

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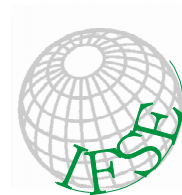


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Worksheets Edited By : Prof. J. Bradley (International Foundation for Science Education)



ADVANCED TEACHING AND LEARNING PACKAGES

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Introductory Note to Teachers

The Microelectrochemistry experiments in this book have been designed for use with the new, upgraded RADMASTE Advanced Microchemistry Kit. Users of the previous version of the RADMASTE Advanced Microchemistry Kit (pre-2006) can still carry out the activities with the older kit, but will need to acquire a few additional items of equipment at low cost.

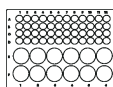
ELECTROLYSIS OF WATER

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet. Tap water is also needed.



2. Equipment

Most of the equipment required can be found in a RADMASTE Advanced Microchemistry Kit. A 9V battery is required. A ruler and permanent marking pen are needed to mark the electrodes into 1 cm units. If the teacher wants the electrolysis to proceed quickly, then the LED can be disconnected from the electrodes and battery after the glowing light has been observed. The electrodes will then need to be connected directly to the battery. This will require extra connecting wires.



3. Hints

A sodium hydroxide pellet is added to the water in the sample vial to increase the conductivity of the water.

The 9V battery is a source of potential difference in the electrical circuit. It must be explained to the learner that the battery does not make the LED glow on its own. If the conducting wires are not placed in some electrolyte solution, the LED will not glow. It may be important to state this so as to avoid the misconception that the battery makes the LED glow, whether a conducting solution is present or not.

If the LED does not glow at first, check that the connections to the battery are secure. It is suggested that connecting wires with crocodile clips be used to ensure that the connections are secure.

The ends of the electrodes in the water must not touch, otherwise the circuit will be completed and the LED will glow brightly, although no electrolysis is occurring. This may be misleading.

Once the current indicator has been disconnected, the continued appearance of tiny bubbles of gas at the electrodes will indicate that a current is still flowing through the water. If no bubbles are evident, check that the connections to the battery and the electrodes are secure.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Sodium hydroxide is a corrosive base. If any base is spilt on the skin, rinse thoroughly with water.

Never point a propette or a syringe containing acid or base upwards. A momentary lapse of concentration can result in a nasty accident. If any acid or base is squirted into the eye, immediately rinse the eye out under running water. In the case of an acid, always have a dilute solution of sodium hydrogencarbonate (household baking soda), or milk close by to apply to the injury. In the case of a base, apply a dilute boric acid solution to the injury. These substances will help neutralise the acid or base in the eye. The patient should be referred to a doctor.

Never allow the learners to play with matches. Treat any burn with cold running water or ice, and seek medical assistance where necessary.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What effect is there on the current indicator when the battery is connected to the electrodes ?

A1. The LED in the current indicator glows.

Q2. What is the reason for your observation in question 1 ?

A2. The LED glows because there is an electric current through the circuit.

Q3. What do you observe at the different electrodes ?

**A3. Electrode 1: Gas bubbles are being generated at the surface of the electrode.
Electrode 2: Gas is also generated at this electrode surface but less than at electrode 1.**

Q4. When electrode 1 is full of substance A, how much of substance B is there in electrode 2 ?

A4. Electrode 2 is approximately half full of substance B when electrode 1 is full of substance A.

Q5. What happens when substance A is exposed to the flame ?

A5. When substance A is exposed to the flame a popping sound can be heard as the gas ignites.

Q6. What is the name given to substance A ?

A6. This explosion is characteristic of gaseous hydrogen (H₂(g)). Thus substance A is hydrogen (H₂(g)).



- Q7. What is the name of substance B ?
- A7. If substance A is gaseous hydrogen ($\text{H}_2(\text{g})$) then the water ($\text{H}_2\text{O}(\text{l})$) must have decomposed. Substance B is therefore probably oxygen ($\text{O}_2(\text{g})$).**
- Q8. What test would you do to prove substance B is what you say it is ?
- A8. The gaseous oxygen ($\text{O}_2(\text{g})$) could be collected until electrode 2 has been filled. A glowing splint could be brought near to the outlet of the electrode. By squeezing the electrode the gas therein could be forced out onto the splint, which should light up brightly as it comes into contact with the gaseous oxygen ($\text{O}_2(\text{g})$).**
- Q9. Why was a greater volume of substance A produced than of substance B ?
- A9. Every molecule of water (H_2O) is composed of two atoms of hydrogen and one atom of oxygen chemically bonded together. Thus, when water (H_2O) is decomposed twice as many hydrogen molecules (H_2) as oxygen molecules (O_2) are generated. This means that the volume of hydrogen should be twice the volume of oxygen.**
- Q10. Write a summary of what happens when water is electrolysed.
- A10. When an electric current is passed through tap water it decomposes into gaseous hydrogen ($\text{H}_2(\text{g})$) and gaseous oxygen ($\text{O}_2(\text{g})$).**
- Q11. From question 10, would you say that tap water is a compound, an element or a mixture ? Explain your answer.
- A11. Water is a compound since when supplied with electrical energy it decomposes into two elements. The proportions in which the two elements, hydrogen and oxygen, are formed is consistent with the formula H_2O for water.**



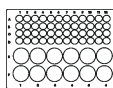
THE ELECTROLYSIS OF A COPPER(II) CHLORIDE SOLUTION

TEACHER GUIDE



1. Chemicals

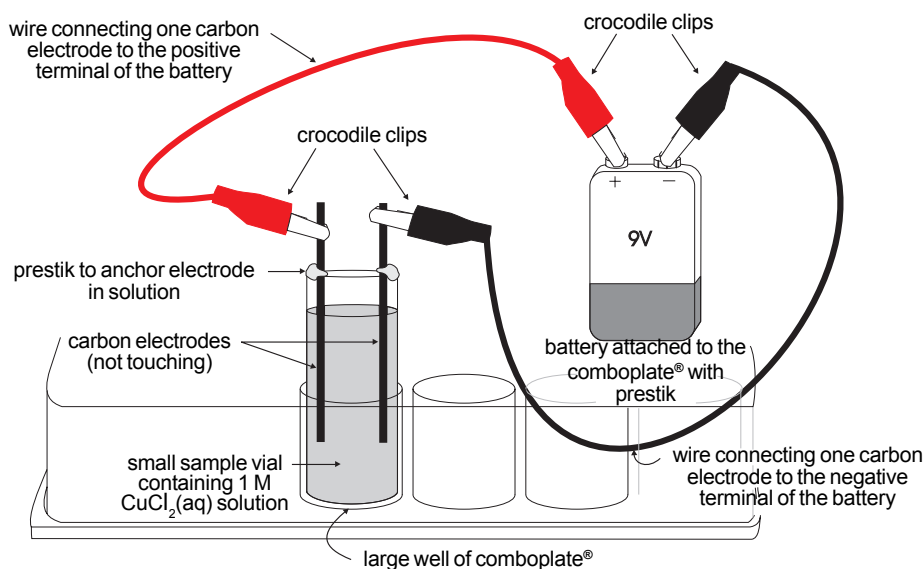
All of the required chemicals are listed in the worksheet. Tap water is also required.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit. Learners will also need a 9V battery, aluminium foil to make two connectors, an ordinary writing ("lead") pencil or two graphite rods, and prestik. Plastic coated paper clips are optional and are used to secure the foil connectors to the ends of each electrode. If connecting wires with crocodile clips are available, these can be used in place of aluminium foil strips to connect the electrodes to the terminals of the battery.

This experiment makes use of a large well of the comboplate® to replace the beaker in the conventional electrolysis of aqueous copper(II) chloride. However, the copper(II) chloride solution may also be electrolysed in the small sample vial in the kit. Users of the Basic Microchemistry Kit can clean and use the sample vial that forms part of the microburner in the kit. Users of the Advanced Microchemistry Kit will already have an extra small sample vial available in the kit, and therefore do not need to use the vial of the microburner. The method that uses the sample vial will require more of the copper(II) chloride solution, but the products of the electrolysis are easier to observe. A diagram of this alternative method appears below.



3. Hints

The 9 V battery is heavy in comparison with the other equipment used in the experiment. It tends to pull the connectors away from the carbon electrodes when it is moved or falls over onto its side. To prevent this, it is suggested that the battery be anchored in a suitable position on the comboplate® with prestik. This position can be altered during the experiment to ensure that the aluminium foil connectors reach the terminals of the battery.

Aluminium foil is an inexpensive material that can be used to construct connectors for the circuit. It is suggested in the procedure that a strip of foil about 3 cm in width and 15 cm in length be used. The strip must then be folded three or four times along the 3 cm edge so that the connector is thick enough to maintain electrical contact without tearing. One end of each foil connector must be attached to a terminal of the battery. Sometimes it is good enough to push the ends of the connectors into the battery terminals. However, it often occurs that the connectors are not in direct contact with the relevant battery terminals and it is therefore recommended that prestik be used to keep the ends of the connectors in place. Similarly, the connectors can slip off the carbon electrodes during the experiment. To prevent this, a plastic-coated paper clip can be gently pushed over the foil where it makes contact with the carbon electrode. If the learners find that electrolysis is not taking place, or that bubbling suddenly stops, they should always check the connections at the battery and at the electrodes to rectify the problem.

Connecting wires with crocodile clips at both ends are also very suitable for maintaining good electrical contact between the electrodes and the battery. Using such wires also enables the learners to identify which electrode is connected to the positive terminal of the battery, and which is connected to the negative terminal. This is necessary for the learners to establish whether an electrode is the cathode or anode, and to determine which products are formed at the different electrodes. The crocodile clips should be as small as possible so that they do not move the electrodes in the solution. Users of the Advanced kits can use the LED because it has connectors with crocodile clips. The bulb also glows when a current is flowing through the solution. However, the LED has a high internal resistance and electrolysis occurs at a much slower rate when the LED forms part of the circuit. It is better to connect the electrodes directly to the battery.



"Lead" or graphite pencils can yield up to four carbon electrodes if the pencil is carefully broken open to remove the central, long carbon rod used for writing. It is important to find pencils that are relatively new or which have not been dropped, otherwise the graphite rod is usually broken at several places within the wooden casing. As an alternative to breaking open a pencil, learners can use the refills for propelling pencils available at stationery stores. The refills need to be approximately 2 mm in diameter so that they do not break when handled. This option is more expensive than breaking open an ordinary pencil. The electrodes should be approximately 5 cm in length to facilitate easy handling during the experiment.

The electrodes must not touch during the experiment as this will result in a completed electrical circuit which excludes the copper(II) chloride solution. If crocodile clips are used they are often heavier than the electrodes and therefore move the electrodes in the solution, sometimes causing the electrodes to touch. Learners should be advised to shift the connectors around until the crocodile clips no longer pull on the electrodes. The well should not be sealed as is sometimes done with the conventional electrolysis of copper(II) chloride. This is because the worksheet allows for the bleaching effect of chlorine to be tested. If the well is sealed, the chlorine gas may not escape from the well and the indicator paper may not be bleached.

The electrolysis begins as soon as the electrodes have been connected to the battery. The production of chlorine gas at the anode is immediate, as shown by vigorous bubbling around the electrode in the solution. The pungent odour of the gas is evident after about one or two minutes from the start of the electrolysis. The bleaching effect of the chlorine on moistened indicator paper can be observed soon after.

The detection of copper metal on the cathode may sometimes be obscured by the deep blue colour of the copper(II) chloride solution. Learners can be encouraged to lift the comboplate® to see the red-brown solid depositing on the electrode. After approximately 5 to 10 minutes, there should be enough copper on the cathode to allow the electrolysis to be stopped. The appearance of the copper deposit varies from scattered granules on the electrode to a fine layer of the metal, which coats the portion of the electrode which has been in contact with the solution. The colour is red-brown in all cases, making for easy identification of the metal.

Early on in the experiment, the learners are asked to identify which electrode is connected to the positive terminal of the battery (Question 3) and which is connected to the negative terminal of the battery (Question 4). Questions 10, 12, 13, 15 and 16 are intended to help learners understand that the electrodes have special names related to the ions in solution which migrate towards them during electrolysis. Learners in the lower grades are not expected to answer the questions on redox reactions.

It is a good idea to allow one or two learners to continue their electrolysis experiments for a longer period of time. They could perhaps leave their comboplate®s overnight in the classroom, or for a few more hours until the school day has ended. Learners will then be able to see a change in the appearance of the copper(II) chloride solution. It is usually pale blue in comparison with the untouched solution in the bottle. The deposit of copper at the cathode is large. Granules of copper fall off the electrode and collect at the bottom of the well.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Chlorine gas is a toxic, pungent gas. The fumes should not be inhaled directly. The experiment should preferably be performed in a ventilated room.

Copper(II) chloride is toxic if ingested. Wash hands thoroughly after the experiment has been completed.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What do you notice as soon as the battery has been connected to the electrodes?

A1. Vigorous bubbling takes place at one of the electrodes.

Q2. Describe the odour coming from the well.

A2. The odour escaping from the well smells like chlorine. (Learners may refer to familiar products like bleach, or the smell associated with swimming pool water.)

Q3. What happens to the section of the indicator paper that is held close to the electrode at which bubbling takes place? Is this electrode connected to the positive or negative terminal of the battery?

**A3. The indicator paper is bleached (i.e it loses its colour and becomes white in appearance.)
The electrode at which bubbling occurs is connected to the positive terminal of the battery.**

Q4. Describe the change in appearance of the other electrode (i.e the electrode where no bubbling occurs). Is this electrode connected to the positive or negative terminal of the battery?

**A4. A red-brown solid is beginning to deposit on the other electrode. (Learners may not be able to detect the colour of the solid early in the experiment as the blue colour of the copper(II) chloride solution sometimes obscures the red-brown colour of the copper deposit.)
This electrode is connected to the negative terminal of the battery.**

- Q5. What has happened to the electrode after the electrolysis of the copper(II) chloride solution has been allowed to continue for 5 to 10 more minutes?
- A5. More of the red-brown solid has deposited on the electrode.**
- Q6. What was happening at the electrode where you saw bubbling taking place? Use your answers to Questions 2 and 3 to support your explanation.
- A6. Bubbles of chlorine gas were formed at the electrode connected to the positive terminal of the battery. This was shown by the smell of chlorine coming from the sample vial, as well as the bleaching effect of the chlorine on the indicator paper.**
- Q7. What was happening at the electrode where no bubbles were observed?
- A7. Copper metal was formed at the electrode where no bubbling took place. This was seen as the red-brown solid coating the electrode.**
- Q8. Describe the appearance of the copper(II) chloride before electrolysis took place. Do the products formed at each electrode have the same properties as the original solution? Explain your answer by referring to observations made during the experiment.
- A8. The copper(II) chloride is a blue solution prior to electrolysis. The products formed at each electrode do not have the same properties as the blue solution. Chlorine is a gas, observed as bubbles during the electrolysis. Copper metal is a solid, seen as red-brown granules on one electrode during and after electrolysis.**
- Q9. From your answer to Question 8, describe the effect of an electric current on a copper(II) chloride solution.
- A9. An electric current causes the copper(II) chloride solution to decompose (or "break down") into copper metal and chlorine gas. (This is an example of a compound being decomposed into its elements by electrolysis.)**
- Q10. The carbon rods or electrodes are required for carrying current into and out of the copper(II) chloride solution. Each electrode has a special name. The electrode connected to the positive terminal of the battery is called the anode, while the electrode connected to the negative terminal of the battery is called the cathode.
- At which electrode did chlorine gas form? (See your answer to Question 3)
 - At which electrode did copper metal deposit? (See your answer to Question 4)
- A10. i. Chlorine gas formed at the anode.
ii. Copper metal deposited at the cathode.**
- Q11. An electric current can only flow if the solution contains charged particles that are able to move through the solution. Write down the formulae of the charged particles which exist in a copper(II) chloride solution. Name the charged particles.
- A11. $\text{Cu}^{2+}(\text{aq})$ and $\text{Cl}^{-}(\text{aq})$. The charged particles or ions are aqueous copper(II) ions and aqueous chloride ions. There are two $\text{Cl}^{-}(\text{aq})$ ions for every one $\text{Cu}^{2+}(\text{aq})$ ion, so the solution is neutral.**
- Q12. Recall what you observed at the anode. Which charged particles in the copper(II) chloride solution moved towards the anode?
- A12. Bubbles of chlorine gas were observed at the anode. This means that the chloride ions in solution moved towards the anode.**
- Q13. Which charged particles moved towards the cathode? Explain by referring to the product you observed at this electrode.
- A13. The copper(II) ions moved towards the cathode because this is the electrode where copper metal was deposited.**
- Q14. Write down a balanced equation to show the reaction taking place in the well during electrolysis. What type of reaction is this? Explain your answer with reference to the observations made at each electrode.
- A14. The reaction taking place during electrolysis can be represented as: $\text{CuCl}_2(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + \text{Cu}(\text{s})$. This is a redox reaction. The oxidation number of the chloride ion in $\text{CuCl}_2(\text{aq})$ is -1. The oxidation number of the chlorine atoms in $\text{Cl}_2(\text{g})$ is 0. Chloride ions have therefore lost electrons and have been oxidised to chlorine atoms. The oxidation number of the copper ion in $\text{CuCl}_2(\text{aq})$ is +2. The oxidation number of the copper atom in $\text{Cu}(\text{s})$ is 0. Copper ions have therefore gained electrons and have been reduced to copper atoms.**
- Q15. What kind of half-reaction occurs at the anode? Write an equation for this half-reaction. (See your answers to Q10i and Q14)
- A15. Oxidation occurs at the anode because chloride ions lost electrons at this electrode and were oxidised to chlorine atoms. The oxidation half reaction occurring at the anode is:
 $2\text{Cl}^{-}(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^{-}$**
- Q16. What kind of half-reaction occurs at the cathode? Write an equation for this half-reaction. (See your answers to Q10ii and Q14)
- A16. Reduction occurs at the cathode because this is the electrode at which copper(II) ions gained electrons and were reduced to copper atoms. The reduction half reaction occurring at the cathode is:
 $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu}(\text{s})$**

CELL POTENTIALS - PART 1

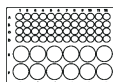
THE ZINC - COPPER CELL

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. A multimeter (or voltmeter) and connecting wires for the multimeter are required. For this experiment it is essential that the multimeter used is accurate.



3. Hints

Users of the older (pre-2006) Advanced Microchemistry Kit can modify the zinc and copper wire coils in the following way to make the zinc and copper electrodes required: Using a set of pliers twist one end of the wire to form two small coils at the one end. The other end should remain straight. The straight end will go into the electrode solution. The electrodes can then be stored as such for future experimentation.

The propette must be thoroughly rinsed before a new solution is dispensed, otherwise the copper sulphate and zinc sulphate solutions will be contaminated.

If the wires from the multimeter are not connected correctly the reading on the multimeter will be negative. The Zn electrode is the anode and must be connected to the negative terminal of the multimeter and the Cu electrode is the cathode and must be connected to the positive terminal of the multimeter. For good results it is very important that the probes of the multimeter are connected properly to the copper and zinc electrodes. It is suggested that connecting wires with crocodile clips be used to connect the multimeter probes securely to the electrodes.

Care should be taken when the copper sulphate and zinc sulphate solutions are added to the respective plastic tips. No air bubbles must be present in the tips containing the solutions. If there is an air bubble it must be removed, otherwise there will be no electrical contact and the multimeter will not record any reading.

It should also be pointed out that the level of the solutions in the two plastic tips should be higher than that of the KCl solution in well F1 to prevent the KCl solution from being pushed up into the plastic tips. As soon as the experiment has been completed, ask the students to remove the cotton wool plugs using the straight ends of the electrodes and washing all the used equipment immediately.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What is the reading (V) on the multimeter ?

A1. The reading should be between 1.02 V and 1.08 V.

Q2. Why is there a reading (V) on the multimeter ?

A2. There is a tendency ("driving force") for the redox reaction to occur in which oxidation takes place at one electrode and reduction takes place at the other. The multimeter reading (the cell potential) measures this tendency.

Q3. At which electrode is oxidation taking place? At which electrode is reduction taking place?

A3. Oxidation occurs at the zinc electrode. Reduction occurs at the copper electrode.

Q4. Which electrode is the anode and which is the cathode ?

A4. The zinc electrode is the anode and the copper electrode is the cathode.

Q5. Write down the half equation to show what is happening at the copper electrode.

A5. Electrons at the electrode surface combine with the Cu^{2+} to produce Cu atoms which stick to the electrode surface. The half equation is : $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$



- Q6. Write down the half equation to show what is happening at the zinc electrode.
- A6. Zn atoms at the electrode surface lose electrons to the external circuit, forming zinc ions. The zinc ions are hydrated and diffuse into solution as $\text{Zn}^{2+}(\text{aq})$. The half equation is : $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$.**
- Q7. Now write down the equation to represent the galvanic cell reaction you have set up.
- A7. $\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$**
- Q8. What is the function of a salt bridge in a galvanic cell ? In this experiment where is the salt bridge ?
- A8. The salt bridge provides electrical connection between the two solutions i.e $\text{CuSO}_4(\text{aq})$ and $\text{ZnSO}_4(\text{aq})$. The KCl solution in the well acts as the salt bridge. (Let the learners see that the galvanic cell set up in this experiment looks upside down compared to traditional textbook pictures.)**
- Q9. What is the standard reduction potential of the copper electrode? What is the standard reduction potential of the zinc electrode?
- A9. This can be read from a table of Standard Electrode Potentials .The standard reduction potential for the copper electrode is 0.34 V. The standard reduction potential for the zinc electrode is - 0.76 V.**
- Q10. What is the standard potential of the zinc - copper cell ?
- A10. The standard cell potential is the sum of the standard reduction potential and the standard oxidation potential of the two half reactions i.e: $0.34 \text{ V} + 0.76 \text{ V} = 1.10 \text{ V}$. (Note : The oxidation potential has the same magnitude but opposite sign to the reduction potential.)**
- Q11. How does the standard cell potential compare with the reading on the multimeter (V) ?
- A11. The standard cell potential should be close to the reading on the multimeter.**

CELL POTENTIALS - PART 2

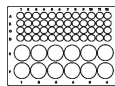
ADDITIVITY

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. A multimeter (or voltmeter) and connecting wires for the multimeter are required. Make sure that the multimeter is reliable.



3. Hints

Users of the older (pre-2006) Advanced Microchemistry Kit can modify the zinc and copper wire coils in the following way to make the zinc and copper electrodes required: Using a set of pliers twist one end of the wire to form two small coils at the one end. The other end should remain straight. The straight end will go into the electrode solution. The electrodes can then be stored as such for future experimentation.

The propette must be thoroughly rinsed before a new solution is dispensed, otherwise the stock solutions will be contaminated. The first part of the experiment is the same as Part 1, i.e : the Zn - Cu cell. You may ask the learners to skip this part if you wish.

It must be explained to the learner that if the wires from the multimeter are not connected correctly the reading on the multimeter will be negative. The Zn electrode is the anode and must be connected to the negative terminal of the multimeter and the Cu electrode is the cathode and must be connected to the positive terminal of the multimeter. Use the same reasoning when the Cu and Pb electrodes are connected and when the Zn and Pb electrodes are connected. For good results it is very important that the probes of the multimeter are connected properly to the zinc, copper and lead electrodes. It is suggested that connecting wires with crocodile clips be used to connect the multimeter probes securely to the electrodes.

Care should be taken when the copper sulphate and zinc sulphate solutions are added to the respective plastic tips. No air bubbles must be present in the tips containing the solutions. If there is an air bubble it must be removed, otherwise there will be no electrical contact and the multimeter will not record any reading.

It should also be pointed out that the level of the solutions in the two plastic tips should be higher than that of the KCl solution in well F1 to prevent the KCl solution from being pushed up into the plastic tips. As soon as the experiment has been completed, ask the students to remove the cotton wool plugs using the straight ends of the electrodes and wash all the equipment immediately.

Note that Questions 11,12 and 13 are suitable for more advanced students, eg. University students.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What is the reading (V) on the multimeter when the copper and zinc electrodes are connected ?

A1. The reading should be between 1.02 V and 1.08 V.

Q2. Write down the equation for the redox reaction in the zinc - copper cell .

A2. $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$

Q3. What is the reading (V) on the multimeter when the copper and lead electrodes are connected ?

A3. The reading (V) on the multimeter should be approximately 0.47 V.

Q4. Write down the equation for the redox reaction in the lead - copper cell.

A4. $\text{Cu}^{2+}(\text{aq}) + \text{Pb(s)} \rightarrow \text{Cu(s)} + \text{Pb}^{2+}(\text{aq})$



- Q5. Predict what will be the reading (V) on the multimeter when the zinc and lead electrodes are connected ? How did you arrive at this answer ? Show all calculations and equations.
- A5. The cell potential for the reaction: $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)} = + 1.10 \text{ V}$**
The cell potential for the reaction: $\text{Cu}^{2+}(\text{aq}) + \text{Pb(s)} \rightarrow \text{Cu(s)} + \text{Pb}^{2+}(\text{aq}) = + 0.47 \text{ V}$
Therefore by subtraction:
The cell potential for the reaction: $\text{Zn(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Pb(s)} = 1.10 \text{ V} - 0.47 \text{ V} = 0.63 \text{ V}$
- Q6. What is the reading (V) on the multimeter when the lead and zinc electrodes are connected ?
- A6. The reading (V) on the multimeter should be approximately 0.63 V.**
- Q7. Write down the equation for the redox reaction in the zinc - lead cell .
- A7. $\text{Zn(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Pb(s)}$**
- Q8. Was your prediction for Question 5 correct?
- A8. If the multimeter is functioning correctly the reading (V) should be close to what you predicted i.e 0.63 V.**
- Q9. What conclusion can you draw from this experiment ?
- A9. The standard potentials of galvanic cells are additive. For this reason if we measure the potentials of two cells in which one electrode is the same, (in this case the Cu/Cu²⁺ electrode) we can use these to predict the potential of the third cell (in this case the Pb-Zn cell) made of the other two electrodes.**
- Q10. If the standard reduction potential of the Cu/Cu²⁺ electrode is + 0.34V , deduce the standard reduction potentials of the Zn/Zn²⁺ and Pb/Pb²⁺ electrodes.
- A10. The measured potential of the Zn/Cu²⁺ reaction should be +1.10 V.**
Therefore $1.10 \text{ V} = \text{Ox. potential of Zn (s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Red. potential of Cu}^{2+}(\text{aq}) \rightarrow \text{Cu(s)}$
Therefore Ox. potential of $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) = 1.10 \text{ V} - 0.34 \text{ V} = 0.76 \text{ V}$
Therefore Red. potential of $\text{Zn}^{2+}(\text{aq}) \rightarrow \text{Zn(s)} = - 0.76 \text{ V}$.
The measured potential of the Cu²⁺/Pb reaction should be 0.47 V.
Therefore $0.47 \text{ V} = \text{Ox. potential of Pb(s)} \rightarrow \text{Pb}^{2+}(\text{aq}) + \text{Red. potential of Cu}^{2+}(\text{aq}) \rightarrow \text{Cu(s)}$
Therefore Ox. potential of $\text{Pb(s)} \rightarrow \text{Pb}^{2+}(\text{aq}) = 0.47 \text{ V} - 0.34 \text{ V} = + 0.13 \text{ V}$
Therefore Red. potential of $\text{Pb}^{2+}(\text{aq}) \rightarrow \text{Pb(s)} = - 0.13 \text{ V}$.
- Q11. For the redox reaction in the Zn-Cu cell, what is the quantity of charge transferred per mole of reaction ?
- A11. For the oxidation of 1 mol of Zn there must be 2 mol e⁻ transferred according to the half equation: $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$. Therefore quantity of charge transferred : $Q_m = 2 \times F = 2 \times (9.649 \times 10^4 \text{ C mol}^{-1}) = 1.930 \times 10^5 \text{ C mol}^{-1}$.**
- Q12. How much work can the Zn-Cu cell perform per mole of reaction ?
- A12. Maximum useful work which the cell can perform : $W_{\text{max}} = Q_m E = (1.930 \times 10^5 \text{ C mol}^{-1}) \times 1.10 \text{ V} = 2.12 \times 10^5 \text{ J mol}^{-1}$ or 212 kJ mol⁻¹.**
- Q13. What is the standard molar free energy for the redox reaction between Zn and Cu²⁺(aq) ?
- A13. The standard molar free energy, $\Delta G^\circ = - W_{\text{max}} = - 212 \text{ kJ mol}^{-1}$.**

CELL POTENTIALS - PART 3

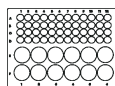
THE EFFECT OF CONCENTRATION

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. A multimeter (or voltmeter) and connecting wires for the multimeter are required. For this experiment it is essential that the multimeter used is accurate. The copper and zinc electrodes used in Part 1 can be used for this experiment as well.



3. Hints

The dilution for the various copper sulphate solutions can be done starting from the 1 M CuSO_4 . If you wish you can do this for the class, or you may ask the learners to do it by themselves. The dilution may be done using tap water. The copper sulphate solution of unknown concentration (it should be between 1 M and 0.001 M) can be prepared by yourself. This experiment can be done in groups of 3 or 4 to complete the experiment more quickly so they have time to answer the questions. The propette must be thoroughly rinsed before a new solution is dispensed, otherwise the solutions will be contaminated.

If the wires from the multimeter are not connected correctly the reading on the multimeter will be negative. The Zn electrode is the anode and must be connected to the negative terminal of the multimeter and the Cu electrode is the cathode and must be connected to the positive terminal of the multimeter. For good results it is very important that the probes of the multimeter are connected properly to the copper and zinc electrodes. It is suggested that connecting wires with crocodile clips be used to connect the multimeter probes securely to the electrodes.

Care should be taken when the copper sulphate and zinc sulphate solutions are added to the respective plastic tips. No air bubbles must be present in the tips containing the solutions. If there is an air bubble it must be removed, otherwise there will be no electrical contact and the multimeter will not record any reading.

It should also be pointed out that the level of the solutions in the two plastic tips should be higher than that of the KCl solution in well F1 to prevent the KCl solution from being pushed up into the plastic tips. As soon as the experiment has been completed, ask the students to remove the cotton wool plugs using the straight ends of the electrodes and wash all the used equipment immediately.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What is the reading (V) on the multimeter? Enter the result in a table like **Table 1** below.

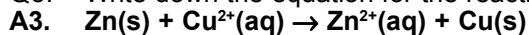
A1. The reading should be approximately 1.10 V (1 100 mV).

Q2. Tabulate the reading on the multimeter (in mV) when the concentration of the copper sulphate solution around the copper electrode is 0.1 M, 0.01 M and 0.001 M respectively.

A2. Table 1

| $[\text{Cu}^{2+}(\text{aq})] / \text{mol dm}^{-3}$ | Log $[\text{Cu}^{2+}(\text{aq})]$ | CELL POTENTIAL / mV |
|--|-----------------------------------|---------------------|
| 1.00 | 0.00 | 1 100 |
| 0.100 | -1.00 | 1 050 |
| 0.0100 | -2.00 | 980 |
| 0.00100 | -3.00 | 880 |

Q3. Write down the equation for the reaction in the Zn - Cu cell.

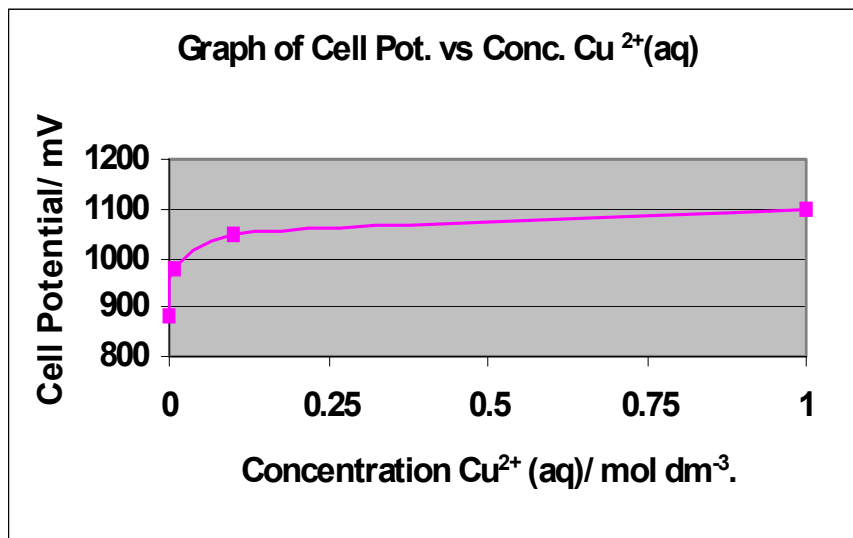


Q4. What did you observe about the reading on the multimeter as you decreased the concentration of the copper sulphate solution ?

A4. **The reading on the multimeter decreased as the concentration of the copper sulphate solution decreased.**

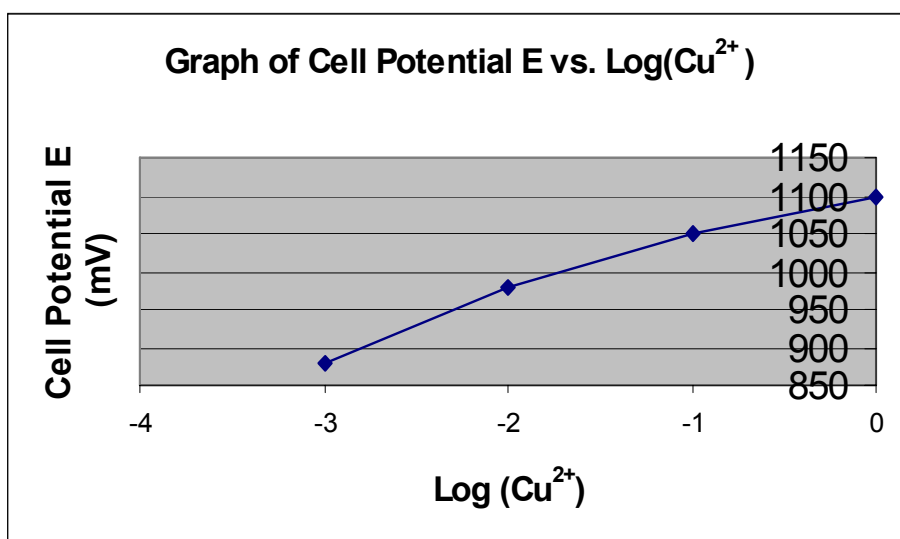
Q5. Draw a graph of cell potential (mV) versus concentration (mol dm^{-3}) of $\text{Cu}^{2+}(\text{aq})$ for the Zn -Cu cell.

A5. **Graph 1 :**



Q6. For each concentration of $\text{Cu}^{2+}(\text{aq})$ calculate the \log_{10} of this concentration and enter the value in Table 1. Draw a graph of cell potential, E versus $\log [\text{Cu}^{2+}(\text{aq})]$ for the Zn - Cu cell. What do you deduce from the graph?

A6. **Graph 2 :**



It can be seen from Graph 2 that Cell Potential is approximately directly proportional to $\log [\text{Cu}^{2+}(\text{aq})]$.

Q7. What is the reading on the multimeter ? What is the concentration of copper ions in the unknown copper sulphate solution ?

A7. **Take the reading from the multimeter (mV) and read off the $\log_{10} [\text{Cu}^{2+}(\text{aq})]$ using Graph 2 . The answer for this question can be compared by letting the learners know what is the concentration of the copper sulphate solution you had prepared. If the experiment was carried out accurately the concentration should compare favourably. You can prepare a solution of different concentration each time you do this experiment OR prepare different solutions for different groups of students in the same class if you wish.**

Q8. Devise a method for measuring the $\text{Zn}^{2+}(\text{aq})$ concentration of a solution.

A8. **This can be done by preparing a series of solutions (from 1 M ZnSO_4) as done above and follow the procedure as for the above experiment but interchanging the CuSO_4 and ZnSO_4 solutions.**

CELL POTENTIALS - PART 4

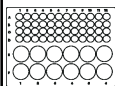
THE NERNST EQUATION

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. A multimeter (or voltmeter) and connecting wires for the multimeter are required. For this experiment it is essential that the voltmeter used is accurate.



3. Hints

The dilution for the various copper sulphate solutions can be done using the 1 M CuSO_4 . If you wish you can do this for the class or you may ask the learners to do it by themselves. N.B: In this part the dilution of the copper sulphate solution must be done using the 1 M sodium sulphate solution instead of tap water. By using this the ionic strength of the solutions is kept constant. This experiment can be done in groups of 3 or 4 to complete the experiment more quickly so the learners have time to answer the questions. The propette must be thoroughly rinsed before a new solution is dispensed, otherwise the solutions will be contaminated.

If the wires from the multimeter are not connected correctly the reading on the multimeter will be negative. The Zn electrode is the anode and must be connected to the negative terminal of the multimeter and the Cu electrode is the cathode and must be connected to the positive terminal of the multimeter. For good results it is very important that the probes of the multimeter are connected properly to the copper and zinc electrodes. It is suggested that connecting wires with crocodile clips be used to connect the multimeter probes securely to the electrodes.

Care should be taken when the copper sulphate and zinc sulphate solutions are added to the respective plastic tips. No air bubbles must be present in the tips containing the solutions. If there is an air bubble it must be removed, otherwise there will be no electrical contact and the multimeter will not record any reading.

It should also be pointed out that the level of the solutions in the two plastic tips should be higher than that of the KCl solution in the vial to prevent the KCl solution from being pushed up into the plastic tips. As soon as the experiment has been completed, ask the students to remove the cottonwool plugs using the straight ends of the electrodes and wash all the used equipment immediately.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required. Note that the Nernst equation is usually only dealt with at the university level.

Q1. What is the reading (V) on the multimeter ? Enter the result in a table like Table 1 below.

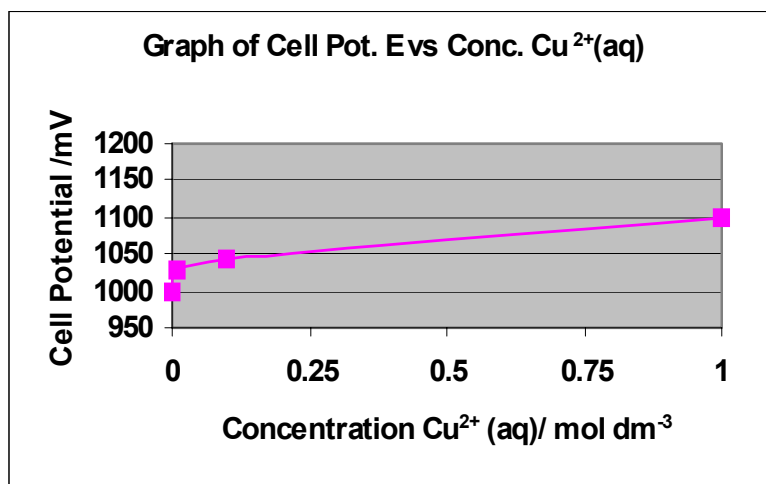
A1. **The reading should be approximately 1.1 V (1 100 mV).**

Q2. Tabulate (Table 1) the reading on the multimeter, ie. the cell potential, E (in mV) when the concentration of the copper sulphate solution around the copper electrode is 0.1 M, 0.01 M and 0.001 M respectively.

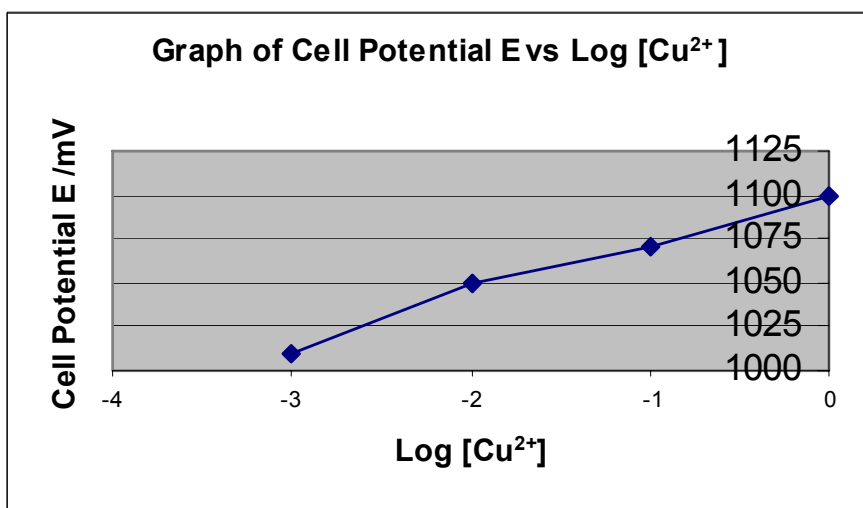
A2. **Table 1:**

| $[\text{Cu}^{2+}(\text{aq})] / \text{mol dm}^{-3}$ | OBSERVED CELL POTENTIAL, E / mV | $\text{Log}[\text{Cu}^{2+}(\text{aq})]$ | $\ln Q$ | EXPECTED CELL POTENTIAL, E / mV |
|--|---------------------------------|---|---------|---------------------------------|
| 1.00 | 1 100 | 0.00 | 0.0 | 1 100 |
| 0.100 | 1 045 | -1.00 | 2.3 | 1 070 |
| 0.0100 | 1 030 | -2.00 | 4.6 | 1 050 |
| 0.00100 | 1 000 | -3.00 | 6.9 | 1 010 |

- Q3. Draw a graph of cell potential E (mV) versus concentration (mol dm^{-3}) of $\text{Cu}^{2+}(\text{aq})$ for the Zn - Cu cell.
 A3. **Graph 1 :**



- Q4. Compare your graph with the one drawn for Question 5 in the previous experiment, ie: Part 3.
 A4. **The graphs look very similar.**
- Q5. For each concentration of $\text{Cu}^{2+}(\text{aq})$ calculate the \log_{10} of this concentration and enter the value in Table 1. Draw a graph of cell potential versus $\log [\text{Cu}^{2+}(\text{aq})]$ for the Zn - Cu cell. Compare your graph with the one drawn for Question 6 in the previous experiment, ie: Part 3.
 A5. **Graph 2 :**
The graphs look very similar.



- Q6. State the Nernst Equation and the numerical value of E^0 , R , T and F under standard conditions.
 A6. **$E = E^0 + (RT/nF) \ln Q$, where**
 E^0 is the standard cell potential (1 100 mV for the Zn - Cu cell)
 n is the number of moles of electrons transferred per mole of reaction (2 in the Zn - Cu cell)
 R is the universal gas constant (8,315 J K⁻¹ mol⁻¹)
 T is the temperature in Kelvin (298 K is the standard)
 F is the Faraday constant (9,650 x 10⁴ C mol⁻¹)
- Q7. What does the quantity Q stand for in the Nernst Equation ?
 A7. **Q is the reaction quotient.**
For our reaction: $\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$
 $[\text{Zn}(\text{s})]$ and $[\text{Cu}(\text{s})]$ are constant.
Hence , $Q = [\text{Zn}^{2+}(\text{aq})] / [\text{Cu}^{2+}(\text{aq})]$
- Q8. Calculate $\ln Q$ for the various copper sulphate solutions and add the values to Table 1.
 A8. **See Table 1 (Remember: For this experiment the concentration of $\text{Zn}^{2+}(\text{aq})$ is 1 M).**
- Q9. Using the Nernst Equation calculate the expected E for each of the $\text{Cu}^{2+}(\text{aq})$ concentrations and enter it in Table 1.
 A9. **See Table 1.**
- Q10. Compare the expected and the experimental values of E .
 A10. **If the experiment is carried out correctly the E values should compare favourably.**

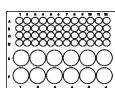
GALVANISING

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. If the 9V battery is not available, use the 2 x 1.5 V cells with a battery holder.



3. Hints

Galvanising will occur at the cathode, so make sure the negative terminal of the battery is connected to the iron electrode (or the carbon electrode). The plastic sample vial is placed in the large well of the comboplate to keep it steady.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What can you observe on the iron electrode ?

A1. A greyish black deposit is seen on the iron electrode.

Q2. Briefly explain your observation in Question 1.

A2. The Zn^{2+} ions in solution at the iron electrode (negative electrode) receive electrons and form zinc atoms. As the zinc atoms accumulate on the electrode surface, they form solid zinc.

Q3. Which electrode is the anode and which is the cathode ?

A3. The zinc electrode is the anode and the iron electrode is the cathode.

Q4. Write down the half - reaction equation to show what happened at the zinc electrode.

A4. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$.

Q5. Write down the half - reaction equation to show what happened at the iron electrode.

A5. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

Q6. What can you observe on the carbon electrode ?

A6. A greyish black deposit, as seen on the iron electrode, is observed on the carbon electrode.

Q7. Rub your finger against the carbon electrode where it was in the zinc sulphate solution. What do you observe?

A7. The greyish black deposit on the carbon electrode comes off onto your finger.

Q8. What is galvanising?

A8. Galvanising is a process, usually electrochemical, in which a coating of zinc is deposited on a solid object (usually made of iron). The zinc coating protects the iron from corrosion (rusting). If the galvanising is done electrochemically it is an example of electroplating.



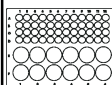
ELECTROPLATING OF COPPER

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. If the 9V battery is not available you may use 2 x 1.5 V cells with a cell holder.



3. Hints

Electroplating will occur at the cathode, so make sure the negative terminal of the battery is connected to the carbon electrode. The plastic sample vial is placed in the large well of the comboplate to make it steady.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What can you observe on the carbon electrode ?

A1. A reddish brown deposit is seen on the carbon electrode.

Q2. Briefly explain your observation in Question 1.

A2. The Cu^{2+} ions in solution at the carbon cathode (negative terminal) receive electrons and form copper atoms. As the copper atoms accumulate on the electrode surface, they form solid copper.

Q3. Which electrode is the anode and which is the cathode ?

A3. The copper electrode is the anode and the carbon electrode is the cathode.

Q4. Write down the half - reaction equation to show what happened at the copper electrode.

A4. $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$

Q5. Write down the half - reaction equation to show what happened at the carbon electrode.

A5. $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$

Q6. What are the requirements for electroplating in terms of the setup ?

A6. For electroplating to occur the metal which is to be electroplated must be the anode and the electrode upon which the electroplating is going to take place must be the cathode.

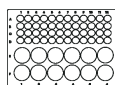
THE OXIDISING POWER OF Fe^{3+} , MnO_4^- and $\text{Cr}_2\text{O}_7^{2-}$

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. A multimeter (or voltmeter) and connecting wires are required. For this experiment it is essential that the multimeter used is accurate.



3. Hints

The propette must be thoroughly rinsed before a new solution is dispensed, otherwise the stock solutions will be contaminated.

It must be explained to the learner that if the wires from the multimeter are not connected correctly the reading on the multimeter will be negative. The Zn electrode is the anode and must be connected to the negative terminal of the multimeter and the carbon electrode is the cathode and must be connected to the positive terminal of the multimeter.

It should also be pointed out that the level of the solutions in the two plastic tips should be higher than that of the KCl solution in well F1 to prevent the KCl solution from being pushed up into the plastic tips. Make sure there are no air bubbles at the bottom of the plastic tips when the solutions are added to it. After the experiment has been performed the cotton wool plugs in the plastic tips can be removed by using the end of the zinc wire electrode. Please remind the learners to wash all used equipment immediately after the experiment has been completed.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What is the reading on the multimeter in V?

A1. The reading should be between 1.48 V and 1.51 V.

Q2. Which electrode is the anode and which is the cathode ?

A2. The zinc electrode is the anode and the carbon electrode is the cathode.

Q3. Write down the half-reaction equation to show what is happening at the zinc electrode.

A3. $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

Q4. Write down the half - reaction equation to show what is happening at the Fe^{3+} , Fe^{2+} / C electrode.

A4. $\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$

Q5. Now write down the equation to represent the galvanic cell reaction you have set up in steps 3 - 6.

A5. $\text{Zn(s)} + 2\text{Fe}^{3+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Fe}^{2+}(\text{aq})$

Q6. Given that the standard reduction potential of the zinc electrode is - 0.76 V, what is the standard reduction potential of the Fe^{3+} , Fe^{2+} / C electrode ?

A6. The standard reduction potential of the Fe^{3+} , Fe^{2+} / C electrode is approximately (1.51 V - 0.76 V) = 0.75 V.

Q7. What is the reading (V) on the multimeter ?

A7. The reading should be between 2.17 V and 2.25 V.

Q8. Write down the half-reaction equation to show what is happening at the H^+ , MnO_4^- , Mn^{2+} / C electrode.

A8. $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)}$

Q9. Write down the equation to represent the galvanic cell reaction you have set up in steps 9 - 10.

A9. $5\text{Zn(s)} + 2\text{MnO}_4^-(\text{aq}) + 16\text{H}^+(\text{aq}) \rightarrow 5\text{Zn}^{2+}(\text{aq}) + 2\text{Mn}^{2+}(\text{aq}) + 8\text{H}_2\text{O(l)}$

- Q10. What is the standard reduction potential of the H^+ , MnO_4^- , Mn^{2+} / C electrode ?
A10. Since the standard reduction potential of the zinc electrode is - 0.76 V , the standard reduction potential of the H^+ , MnO_4^- , Mn^{2+} / C electrode is approximately $(2.25 \text{ V} - 0.76 \text{ V}) = 1.49 \text{ V}$.
- Q11. What is the reading (V) on the multimeter ?
A11. The reading should be between 2.06 V and 2.10 V.
- Q12. Write down the half-reaction equation to show what is happening at the H^+ , $\text{Cr}_2\text{O}_7^{2-}$, Cr^{3+} / C electrode.
A12. $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$
- Q13. Write down the equation to represent the galvanic cell reaction you have set up in step 10.
A13. $3\text{Zn}(\text{s}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) \rightarrow 3\text{Zn}^{2+}(\text{aq}) + 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$
- Q14. What is the standard reduction potential of the H^+ , $\text{Cr}_2\text{O}_7^{2-}$, Cr^{3+} / C electrode ?
A14. Since the standard reduction potential of the zinc electrode is - 0.76 V , the standard reduction potential of the $\text{Cr}_2\text{O}_7^{2-}$, Cr^{3+} / C electrode is approximately $(2.10 \text{ V} - 0.76 \text{ V}) = 1.34 \text{ V}$.
- Q15. Fe^{3+} , $\text{Cr}_2\text{O}_7^{2-}$ and MnO_4^- are all oxidising agents. Rank them in order from the strongest to the weakest oxidising agent .
A15. The strongest oxidising agent is MnO_4^- followed by $\text{Cr}_2\text{O}_7^{2-}$, and the weakest oxidising agent is Fe^{3+} .
- Q16. What is the function of the carbon electrode in this experiment ?
A16. The carbon electrode acts merely as an electrical conductor in this experiment .

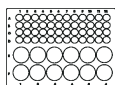
THE HYDROGEN - OXYGEN FUEL CELL

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit. A multimeter and connecting wires for the multimeter are required. If the 9V battery is not available you may use 2 x 1.5 V cells with a cell holder.



3. Hints

This experiment is similar to the experiment on the electrolysis of water.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Sodium hydroxide is a corrosive base. If any base is spilt on the skin, rinse thoroughly with water.

Never point a propette or a syringe containing acid or base upwards. A momentary lapse of concentration can result in a nasty accident. If any acid or base is squirted into the eye, immediately rinse the eye out under running water. In the case of an acid, always have a dilute solution of sodium hydrogencarbonate (household baking soda), or milk close by to apply to the injury. In the case of a base, apply a dilute boric acid solution to the injury. These substances will help neutralise the acid or base in the eye. The patient should be referred to a doctor.

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

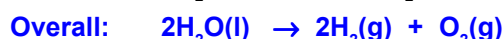
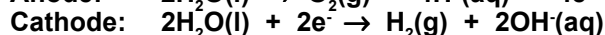
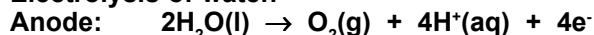
It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What do you observe when the battery is connected to the carbon electrodes ?

A1. Bubbles of gas are produced at both the carbon electrodes.

Q2. What do you call the chemical reaction that is observed in Question 1? Write the chemical equations for the reactions that occur at the anode, the cathode and for the overall reaction.

A2. Electrolysis of water.



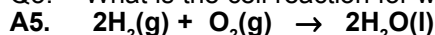
Q3. Which substance is produced in the carbon electrode that fills up first? Give a reason for your answer.

A3. Hydrogen gas ($\text{H}_2(\text{g})$) is produced in the carbon electrode that fills up first. When a water molecule breaks down there are twice the number of hydrogen atoms compared to the oxygen atoms. Therefore twice as many hydrogen molecules, H_2 , as oxygen molecules, O_2 , are produced.

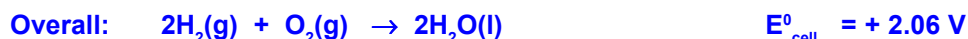
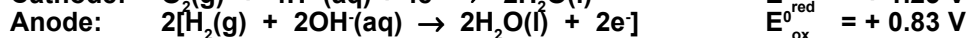
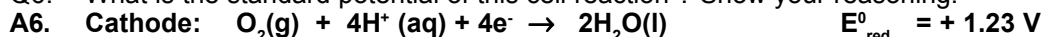
Q4. What is the reading on the multimeter ?

A4. The reading should be approximately 1.68 V.

Q5. What is the cell reaction for which the potential has been measured ?



Q6. What is the standard potential of this cell reaction? Show your reasoning.



Q 7. How does your answer in Question 6 compare with the reading on the multimeter ?

A 7. Theoretically, the reading on the multimeter should be close to 2.06 V, but experimentally the maximum voltage obtained is about 1.7 V. The difference is largely due to the non-standard concentrations and pressures used in the experiment.



Q8. What is a fuel cell ? How can the H_2/O_2 cell function as a fuel cell ?

A8. A fuel cell is a galvanic cell in which the reactants are continuously supplied and the products are continuously removed as current is drawn from the cell. To use the cell we have made as a fuel cell we would need to supply more oxygen and hydrogen from an outside source as we used the cell. The carbon electrodes work well because they absorb small quantities of gaseous hydrogen and oxygen. The hydrogen and oxygen are the fuels for an energy - releasing reaction occurring at the carbon electrodes in the two half - reactions. As the reaction proceeds , water is produced, and we would need to keep removing it (or allow it to over-flow).



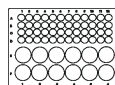
USING CONDUCTIVITY TO DISTINGUISH BETWEEN METALS AND NON METALS

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Apart from the bar LED conductivity indicator, most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit.

The bar LED conductivity indicator measures conductivity by displaying a series of lit bars (in the upper right corner of the comboplate®), when the electrodes of the instrument are in contact with a conductor. It requires a 9V battery to function.



3. Hints

When using the electrodes to measure the conductivity of solid substances like the metal strips or iron nail, the electrodes must be flat against the substance being tested.

When measuring the conductivity of powders, like sulphur powder, make sure the electrodes are held gently below the surface of the powder. **DO NOT PUSH THE ELECTRODES AGAINST THE BOTTOM OF THE WELL**, as this can result in damage to the wires.

Solid crystals should be large enough that both electrodes are touching the solid.

Wait a few seconds before recording the number of bars lit, as there may be some fluctuation in the conductivity when the electrodes are first placed in contact with the sample.

Do not allow the electrodes to touch the sides of the well, as this will influence the conductivity reading.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.

Iodine vapour is toxic; make sure the room is well ventilated.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. Prepare a table like Table 1 below and record your observations there.

A1. TABLE 1: THE CONDUCTIVITY OF SOME METALS AND NON METALS USING THE BAR LED CONDUCTIVITY INDICATOR

| MATERIAL TESTED | NUMBER OF BARS GLOWING |
|-----------------|------------------------|
| Magnesium | 10 |
| Lead | 10 |
| Iron | 10 |
| Copper | 10 |
| Nickel | 10 |
| Zinc | 10 |
| Aluminium | 10 |
| Sulphur | 0 |
| Iodine | 0 |
| Carbon | 10 |

- Q2. What do you notice about all the substances tested?
(Hint: are they elements, compounds or mixtures?)
A2. They are all elements.
- Q3. Which substance/s tested conduct electricity?
A3. Magnesium, lead, iron, copper, nickel, zinc, aluminium and carbon.
- Q4. Which substance/s tested do not conduct electricity?
A4. Iodine and sulphur.
- Q5. What is the difference between the conductors and non-conductors of electricity?
A5. Most of the conductors are metals. The non-conductors are non metals.
- Q6. What is special about carbon (graphite)?
A6. It is a non metal, yet it has electrical conductivity.

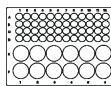
USING CONDUCTIVITY TO DISTINGUISH BETWEEN IONIC AND COVALENT COMPOUNDS

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet. Deionised water is required to dissolve the solid compounds as well as to rinse the electrodes between each use.



2. Equipment

Apart from the bar LED conductivity indicator, most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit. The paper cup (or any other suitable container) is used to catch the waste from the rinsing of the electrodes.

The bar LED conductivity indicator measures conductivity by displaying a series of lit bars (in the upper right corner of the comboplate®), when the electrodes of the instrument are in contact with a conductor or immersed in a conducting solution. It requires a 9V battery to function.



3. Hints

In this activity, learners observe that covalent compounds do not conduct electricity, whether in solid form or as a solution. They also learn that ionic compounds in solid form do not conduct an electric current, but when the solids are dissolved in deionised water the resulting solutions do conduct electricity. To be able to successfully show these differences in conductivity, the wells of the comboplate® must be thoroughly cleaned. Any contamination in the wells, such as residue from a previous experiment, can dissolve in the deionised water added during later steps to the solid compounds. This can result in confusing conductivity results being obtained. For example, a sucrose solution should not cause any bars to light up on the multimeter. If one or more bars begins to glow, this indicates that some ionic compound may have been present in the well to which the solid sucrose was added. The ionic compound also dissolves in the deionised water, resulting in a solution that conducts an electric current. Similarly, the solid compounds to be tested should be pure, otherwise impurities in the solid samples can also influence conductivity when they dissolve in the deionised water.

When using the microspatula to add a solid to a large well, care must be taken not to allow any of the compound to spill into an adjacent well. If this happens, the adjacent well should not be used. Rather skip the well in which spillage has occurred and use the next, clean well for testing a different solid sample.

When measuring the conductivity of solid samples, make sure that the electrodes are held gently below the surface of the solid. **DO NOT PUSH THE ELECTRODES AGAINST THE BOTTOM OF THE WELL**, as this can result in damage to the wires.

It is very important to remove the electrode probe from the solution. The bar LED conductivity indicator has a DC (Direct Current) circuit. If the probe is left in the conducting solution, electrolysis will occur and affect the conductivity measured. If any droplets of solution remain on the electrode wires, the same problem will occur. It is therefore also important to transfer any solution from the wires into the bulk of the solution using the microspatula, after recording the conductivity measured.

Do not allow the electrodes to touch the sides of the well, as this will influence the conductivity reading.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. Prepare a table like Table 1 below and record your observations there.

A1. TABLE 1: THE CONDUCTIVITY OF SOME SOLIDS AND THEIR SOLUTIONS USING THE BAR LED CONDUCTIVITY INDICATOR

| SUBSTANCE TESTED | CONDUCTIVITY/NUMBER OF BARS GLOWING |
|--|-------------------------------------|
| $\text{KNO}_3(\text{s})$ | 0 |
| $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$ | 0 |
| $\text{NaCl}(\text{s})$ | 0 |
| $\text{KNO}_3(\text{aq})$ | 10 |
| $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{aq})$ | 0 |
| $\text{NaCl}(\text{aq})$ | 9 or 10 |

Q2. Which of the solids conducts electricity?

A2. None of the solids conduct electricity.

Q3. Which solution/s tested conducts electricity?

A3. The $\text{NaCl}(\text{aq})$ and $\text{KNO}_3(\text{aq})$ solutions conduct electricity.

Q4. Which solution/s do not conduct electricity?

A4. The sucrose solution does not conduct electricity.

Q5. Why do some solutions conduct electricity and some do not?

A5. A solution will only conduct electricity if it contains mobile ions that carry charge. Those solutions that contain ions conduct electricity while those without ions do not.

Q6. Why do the solutions $\text{NaCl}(\text{aq})$ and $\text{KNO}_3(\text{aq})$ conduct electricity whereas their crystals do not?

A6. In $\text{NaCl}(\text{s})$ and $\text{KNO}_3(\text{s})$ there may be ions. However they cannot move as they are tightly bonded to each other in an infinite 3-D lattice structure. When these solids dissolve in water, separate ions are formed which are hydrated (i.e they have several water molecules bonded to them). Therefore the solids do not conduct electricity whereas their solutions do.

Q7. Why do both the sucrose crystals and the sucrose solution fail to conduct an electric current?

A7. In the solid sucrose, each molecule of sucrose is packed closely to other sucrose molecules. There are no ions. When the sucrose crystal is dissolved, intermolecular forces are broken and entire sucrose molecules are released into solution. As a result, no current can be conducted because the molecules are neutral and will not carry charge.

Q8. Which of the compounds tested, would you describe as ionic and which as covalent?

A8. Sodium chloride and potassium nitrate are described as ionic compounds. Sucrose is described as a covalent compound.

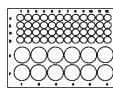
THE CONDUCTIVITY OF SOME LIQUIDS

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet. Deionised water is required to dissolve the solid compounds as well as to rinse the electrodes between each use.



2. Equipment

Apart from the bar LED conductivity indicator, most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit. The paper cup (or any other suitable container) is used to catch the waste from the rinsing of the electrodes.

The bar LED conductivity indicator measures conductivity by displaying a series of lit bars (in the upper right corner of the comboplate®), when the electrodes of the instrument are in contact with a conductor or immersed in a conducting solution. It requires a 9V battery to function.



3. Hints

The wells of the comboplate® which are used for testing, should be rinsed with deionised water and dried thoroughly. If the wells contain contaminants, these may dissolve in the liquids you are testing and affect the conductivity obtained.

The teacher must ensure that contamination of the test liquids is prevented, because they are all clear. This can be done by either labelling each propette with the name of a particular liquid, or by placing the propettes in specially demarcated areas immediately after chemicals of concern have been dispensed.

Wait a few seconds before recording the number of bars lit, as there may be some fluctuation in the conductivity when the electrodes are first placed into the liquid.

DO NOT PUSH THE ELECTRODES AGAINST THE BOTTOM OF THE WELL, as this can result in damage to the wires.

It is very important to remove the electrode probe from the liquid. The bar LED conductivity indicator has a DC (Direct Current) circuit. If the probe is left in a conducting liquid, electrolysis will occur and affect the conductivity measured. If any droplets of liquid remain on the electrode wires, the same problem will occur. It is therefore also important to transfer any liquid from the wires into the bulk of the liquid using the microspatula, after recording the conductivity measured.

Do not allow the electrodes to touch the sides of the well, as this will influence the conductivity reading.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.

The organic reagents cyclohexane, cyclohexene and ethanol form toxic fumes. Avoid inhaling these fumes and make sure the experiment is performed in a well ventilated room.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. Prepare a table like Table 1 below and record your observations there.

A1. TABLE 1: THE CONDUCTIVITY OF SOME LIQUIDS USING THE BAR LED CONDUCTIVITY INDICATOR

| LIQUID TESTED | CONDUCTIVITY/NUMBER OF BARS GLOWING |
|---|-------------------------------------|
| Tap water (H ₂ O(l)) | 3 |
| Distilled water (H ₂ O(l)) | 1 |
| Deionised water (H ₂ O(l)) | 0 |
| C ₆ H ₁₂ (l) | 0 |
| C ₆ H ₁₀ (l) | 0 |
| C ₂ H ₅ OH(l) | 0 |
| HOCH ₂ CH(OH)CH ₂ OH(l) | 0 |

Q2. Which liquid/s tested are electrical conductors?

A2. Tap water and distilled water.

Q3. Why do the three samples of water have different conductivities?

A3. Tap water contains ions and other chemicals introduced during water purification. The ions cause the water to conduct a current. Distilled water has had organic and other impurities removed. Some ions are present in the distilled water but the concentration of these ions is less than in tap water, resulting in the lower conductivity measured. Deionised water has had all impurity ions removed and the very small concentration (10⁻⁷ M) of OH⁻(aq) and H₃O⁺(aq) ions is too low for a current to be detected.

Q4. Why do the pure liquids not conduct an electric current?

A4. These compounds consist only of molecules. There are no ions available for the transport of electrical charge.

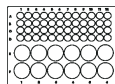
THE EFFECT OF ELECTROLYTE CONCENTRATION ON CONDUCTIVITY

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet. Deionised water is required to rinse the electrodes between measurements and at the end of the activity.



2. Equipment

Apart from the bar LED conductivity indicator, most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit. The paper cup (or any other suitable container) is used to catch the waste from the rinsing of the electrodes.

The bar LED conductivity indicator measures conductivity by displaying a series of lit bars (in the upper right corner of the comboplate®), when the electrodes of the instrument are in contact with a conductor or immersed in a conducting solution. It requires a 9V battery to function.



3. Hints

The learners may use their own comboplate®s for this activity. The electrode probe from the bar LED conductivity meter is long enough to allow the electrodes to be immersed into a solution/s in the wells of another comboplate® placed alongside. Each learner must rinse and dry the electrodes before the next learner uses the meter.

Wait a few seconds before recording the number of bars lit, as there may be some fluctuation in the conductivity when the electrodes are first placed into the solution.

DO NOT PUSH THE ELECTRODES AGAINST THE BOTTOM OF THE WELL, as this can result in damage to the wires.

It is very important to remove the electrode probe from the solution. The bar LED conductivity indicator has a DC (Direct Current) circuit. If the probe is left in the conducting solution, electrolysis will occur and affect the conductivity measured. If any droplets of solution remain on the electrode wires, the same problem will occur. It is therefore also important to transfer any solution from the wires into the bulk of the solution using the microspatula, after recording the conductivity measured.

Do not allow the electrodes to touch the sides of the well, as this will influence the conductivity reading.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.

The organic reagents cyclohexane, cyclohexene and ethanol form toxic fumes. Avoid inhaling these fumes and make sure the experiment is performed in a well ventilated room.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. Prepare a table like Table 1 below and record your observations there.

A1. **TABLE 1: THE EFFECT OF ELECTROLYTE CONCENTRATION ON CONDUCTIVITY**

| CONCENTRATION OF NaCl(aq) TESTED/ M | CONDUCTIVITY/NUMBER OF BARS GLOWING |
|-------------------------------------|-------------------------------------|
| 1×10^{-3} | 3 |
| 0.01 | 6 |
| 0.05 | 8 |

- Q2. What happens to the electrical conductivity of NaCl(aq) as its concentration is increased? Explain your answer.
- A2. The conductivity increases. The greater the concentration of the NaCl(aq) solution, the greater the number of Na⁺(aq) and Cl⁻(aq) ions there are per volume. More ions per volume means more charge carriers per volume and the more concentrated NaCl(aq) solutions therefore cause more bars to glow.**
- Q3. Make a general statement about the relationship between conductivity and electrolyte concentration.
- A3. Conductivity increases as electrolyte concentration increases.**

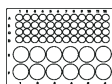
USING CONDUCTIVITY TO DISTINGUISH BETWEEN STRONG AND WEAK ACIDS AND BASES

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet. Deionised water is required to rinse the electrodes once the activity has been completed.



2. Equipment

Apart from the bar LED conductivity indicator, most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit. The paper cup (or any other suitable container) is used to catch the waste from the rinsing of the electrodes.

The bar LED conductivity indicator measures conductivity by displaying a series of lit bars (in the upper right corner of the comboplate®), when the electrodes of the instrument are in contact with a conductor or immersed in a conducting solution. It requires a 9V battery to function.



3. Hints

The learners may use their own comboplate®s for this activity. The electrode probe from the bar LED conductivity meter is long enough to allow the electrodes to be immersed into a solution/s in the wells of another comboplate® placed alongside. Each learner must rinse and dry the electrodes before the next learner uses the meter.

The teacher must ensure that contamination of the test liquids is prevented, because they are all clear. This can be done by either labelling each propette with the name of a particular solution, or by placing the propettes in specially demarcated areas immediately after chemicals of concern have been dispensed.

Wait a few seconds before recording the number of bars lit, as there may be some fluctuation in the conductivity when the electrodes are first placed into the liquid.

DO NOT PUSH THE ELECTRODES AGAINST THE BOTTOM OF THE WELL, as this can result in damage to the wires.

It is very important to remove the electrode probe from the solution. The bar LED conductivity indicator has a DC (Direct Current) circuit. If the probe is left in the conducting solution, electrolysis will occur and affect the conductivity measured. If any droplets of solution remain on the electrode wires, the same problem will occur. It is therefore also important to transfer any solution from the wires into the bulk of the solution using the microspatula, after recording the conductivity measured.

Do not allow the electrodes to touch the sides of the well, as this will influence the conductivity reading.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Hydrochloric and acetic acids are corrosive. If any acid is spilt on the skin, the affected area must immediately be rinsed with copious amounts of water. Severe burns must receive medical attention.

Sodium hydroxide and ammonia are corrosive bases. If any base is spilt on the skin, treat as above for acid burns.

Never point a propette or a syringe containing acid or base upwards. A momentary lapse of concentration can result in a nasty accident. If any acid or base is squirted into the eye, immediately rinse the eye out under running water. In the case of an acid, always have a dilute solution of sodium hydrogencarbonate (household baking soda), or milk close by to apply to the injury. These substances will help neutralise the acid in the eye. The patient should be referred to a doctor.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. Prepare a table like Table 1 below and record your observations there.

A1. TABLE 1: CONDUCTIVITY OF SOME STRONG AND WEAK ACIDS AND BASES

| ACID/BASE | CONDUCTIVITY/NUMBER OF BARS GLOWING |
|--------------------------------|-------------------------------------|
| 0.1 M HCl(aq) | 10 |
| 0.1 M CH ₃ COOH(aq) | 6 |
| 1 M CH ₃ COOH(aq) | 8 |
| 0.1 M NaOH(aq) | 6 |
| 0.1 M NH ₃ (aq) | 3 |
| 1 M NH ₃ (aq) | 4 |

Q2. Which acid conducts electricity better, the 0.1 M HCl(aq) or the 0.1 M CH₃COOH(aq)?

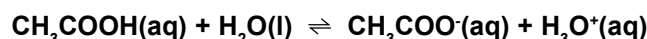
A2. The 0.1 M HCl(aq) conducts electricity better.

Q3. Which base is the better electrical conductor, 0.1 M NaOH(aq) or 0.1 M NH₃(aq)?

A3. The 0.1 M NaOH(aq) is the better electrical conductor.

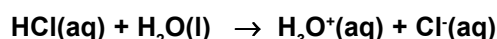
Q4. A solution conducts an electric current because it contains mobile particles (ions) that carry charge. The greater the concentration of ions in solution, the higher the solution conductivity. Since one molecule of HCl and one molecule of CH₃COOH can each yield one negative and one positive ion in solution and their solutions are identical in concentration, why do they have such different conductivities?

A4. Although the acid solutions have the same concentration, the lower conductivity of the CH₃COOH(aq) suggests that there is a lower concentration of ions in this solution than in the HCl(aq). This is because the CH₃COOH molecules do not react completely when mixed with water. The equation representing this is:



Every CH₃COOH molecule therefore does not yield one positive and one negative ion because the reaction with water is not complete. At any one time in solution there exists some CH₃COO⁻(aq) ions, some H₃O⁺(aq) ions and some CH₃COOH(aq) molecules. The ions conduct current but the molecules do not.

In contrast, every HCl molecule reacts with water to form H₃O⁺(aq) and Cl⁻(aq) ions:



There are proportionately more ions per volume in the solution of 0.1 M HCl(aq) than in 0.1 M CH₃COOH(aq), resulting in the higher HCl(aq) conductivity.

Q5. Based on your answer to Question 4, which of the acids is weak and which is strong?

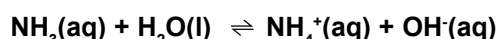
A5. The HCl is strong and the CH₃COOH is weak.

Q6. What further evidence is there from the experiment to support your answer to Question 5?

A6. The 1 M CH₃COOH(aq) has a lower conductivity than the 0.1 M HCl(aq) even though it is ten times greater in concentration. This proves that the HCl(aq) contains a greater concentration of conducting ions than the CH₃COOH(aq), because the latter is only partially ionised in solution.

Q7. Which of the bases tested is stronger, NaOH or NH₃? Explain your answer.

A7. The NaOH is the stronger base as shown by its higher conductivity. When NaOH reacts with water, it is completely dissociated into Na⁺(aq) and OH⁻(aq) ions. In contrast, NH₃ does not react completely with water:



The 0.1 M NaOH(aq) has more ions per volume than both the 0.1 M and 1 M NH₃(aq) solutions, and therefore a higher conductivity.

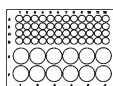
USING CONDUCTIVITY TO DISTINGUISH BETWEEN COMPLETELY DISSOCIATED AND PARTIALLY DISSOCIATED SALTS

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet. Deionised water is required to rinse the electrodes between different test solutions, and once the activity has been completed.



2. Equipment

Apart from the bar LED conductivity indicator, most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit. The paper cup (or any other suitable container) is used to catch the waste from the rinsing of the electrodes.

The bar LED conductivity indicator measures conductivity by displaying a series of lit bars (in the upper right corner of the comboplate®), when the electrodes of the instrument are in contact with a conductor or immersed in a conducting solution. It requires a 9V battery to function.



3. Hints

The learners may use their own comboplate®s for this activity. The electrode probe from the bar LED conductivity meter is long enough to allow the electrodes to be immersed into a solution/s in the wells of another comboplate® placed alongside. Each learner must rinse and dry the electrodes before the next learner uses the meter.

The teacher must ensure that contamination of the test liquids is prevented, because they are both clear. This can be done by either labelling each propette with the name of a particular solution, or by placing the propettes in specially demarcated areas immediately after chemicals of concern have been dispensed.

Wait a few seconds before recording the number of bars lit with the NaCl(aq), as there may be some fluctuation in the conductivity when the electrodes are first placed into the solution. However, since mercury(II) chloride rapidly corrodes the electrodes, learners must remove the probe promptly from this solution.

DO NOT PUSH THE ELECTRODES AGAINST THE BOTTOM OF THE WELL, as this can result in damage to the wires.

It is very important to remove the electrode probe from the solution. The bar LED conductivity indicator has a DC (Direct Current) circuit. If the probe is left in the conducting solution, electrolysis will occur and affect the conductivity measured. If any droplets of solution remain on the electrode wires, the same problem will occur. It is therefore also important to transfer any solution from the wires into the bulk of the solution using the microspatula, after recording the conductivity measured.

Do not allow the electrodes to touch the sides of the well, as this will influence the conductivity reading.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin. Mercury(II) chloride is poisonous, so please ask the learners to be careful when handling it.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. Prepare a table like Table 1 below and record your observations there.

A1. TABLE 1: THE CONDUCTIVITY OF COMPLETELY DISSOCIATED AND PARTIALLY DISSOCIATED SALTS

| SOLUTION | CONDUCTIVITY/NUMBER OF BARS GLOWING |
|------------------------------|-------------------------------------|
| 0.1 M NaCl(aq) | 8 |
| 0.1 M HgCl ₂ (aq) | 6 |

Q2. Which salt solution is the better electrolyte?

A2. The NaCl(aq) is the better electrolyte.

Q3. When a crystal of NaCl(s) dissolves in water, the ionic bonds are broken forming Na⁺(aq) and Cl⁻(aq) ions: NaCl(s) → Na⁺(aq) + Cl⁻(aq). When a crystal of HgCl₂(s) is placed in water, HgCl₂(s) molecules mix with water molecules and some of them dissociate according to the equation:



Use this information to explain your answer to Question 2 and the conductivity results obtained.

A3. The ionic NaCl(s) crystals react completely with water so that the NaCl(aq) consists entirely of Na⁺(aq) and Cl⁻(aq) ions. The covalent HgCl₂(s) crystals consist of discrete molecules that only partially form ions in water. As a result, at any one time in solution there exists some HgCl₂(aq) molecules, some HgCl⁺(aq) ions and some Cl⁻(aq) ions. The lower concentration of ions in HgCl₂(aq) results in a solution with lower conductivity.

ASSEMBLY AND USE OF THE MICROBURETTE IN MICROSCALE VOLUMETRIC ANALYSIS

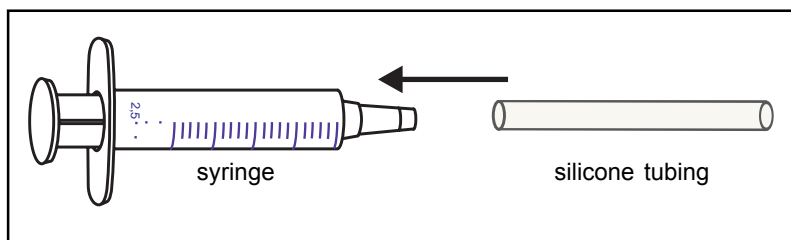
INTRODUCTORY NOTES

Assembling the Microburette

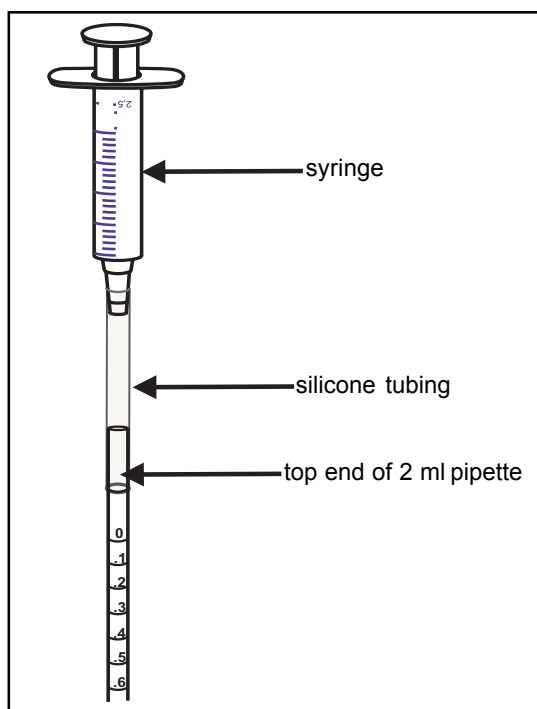
APPARATUS

- ♣ 2 x 2 ml plastic syringes
- ♣ 2 x 2 ml plastic pipettes graduated in 0.01 ml intervals
- ♣ 2 x plastic pipette tips
- ♣ 2 x silicone tubes (~ 4 cm in length)
- ♣ 2 x plastic microstands
- ♣ 1 x comboplate®
- ♣ 2 x propettes
- ♣ microspatulas

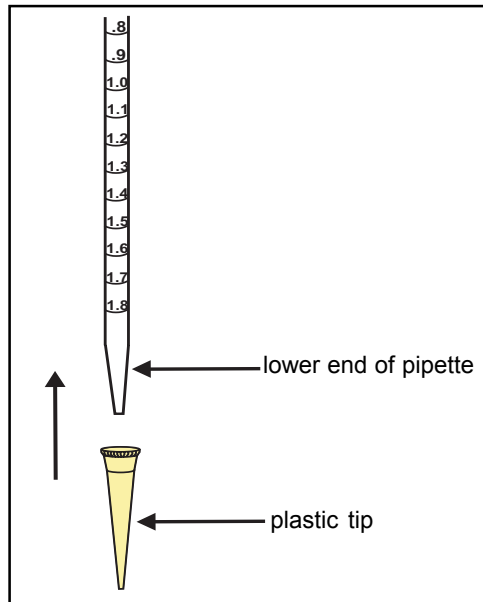
1. Attach one end of the silicone tubing to the 2 ml syringe by gently sliding the tubing over the nozzle of the syringe.



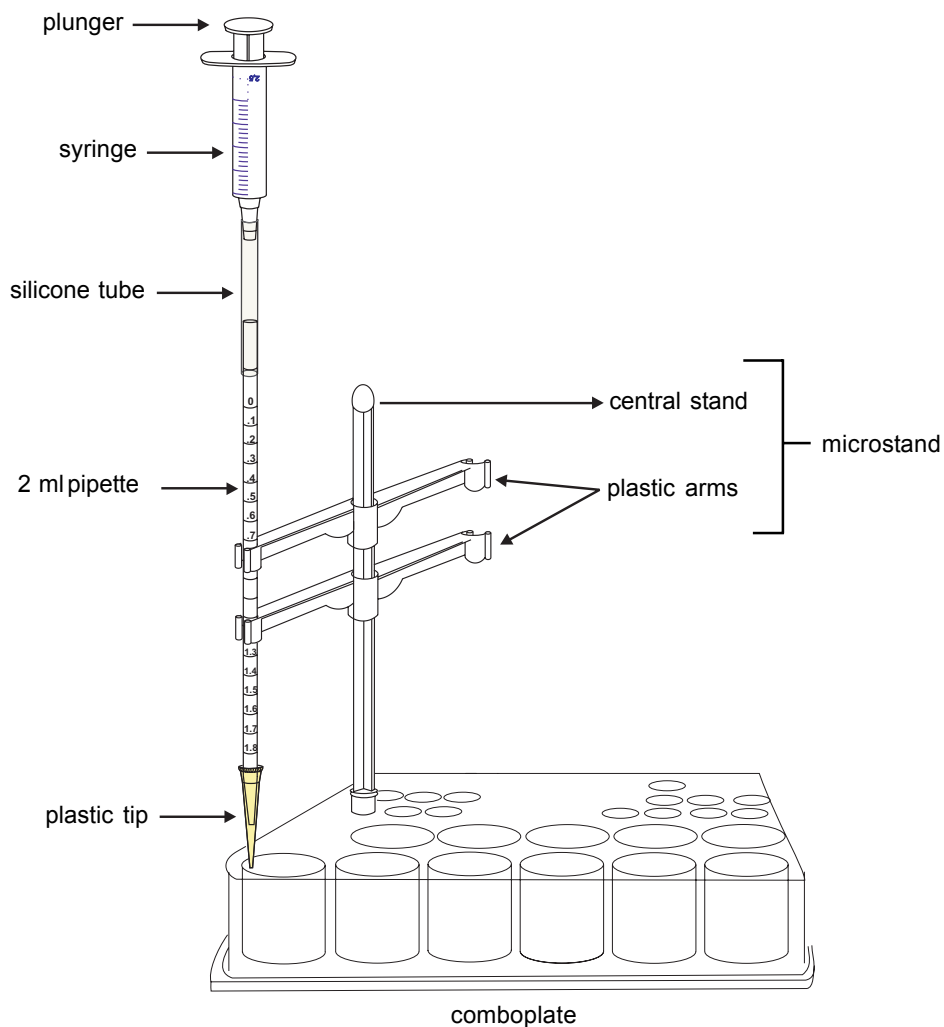
2. Connect the remaining end of the silicone tubing to the 2 ml pipette by similarly sliding the tubing over the top of the pipette.



3. Place the plastic tip on the bottom, pointed end of the pipette. Make sure that the tip is secure to prevent it from falling off during the titration procedure.



4. Assemble the microstand by placing two plastic arms onto one of the central retorts. Clamp the now-assembled microburette in an upright position by clipping the pipette into each arm of the microstand.
5. Push the microstand into one of the small wells in the comboplate®. Adjust the position of the microburette by sliding the pipette up or down the arms of the microstand.



Operating the Microburette

The following instructions are intended as a practice session for operating the microburette. Water is used to represent the titrant solution. For this reason, excess water in the microburette can be returned to the water container. However, remember that when an actual titration is performed excess solution should never be returned to the original titrant container, but discarded as waste.

REQUIREMENTS

- ♣ 1 x microburette
- ♣ 1 x waste container
- ♣ 1 x comboplate®
- ♣ 1 x microspatula
- ♣ water

A. Rinsing the microburette

1. Slide down the plunger of the syringe until it is all the way inside the syringe. The plunger should move easily up and down inside the syringe. If it is tight, you may need to pump the plunger inside the barrel of the syringe until it slides smoothly. Alternatively, the plunger can be replaced with another that slides easily.
2. Hold the container of water (or titrant solution) under the microburette so that the plastic tip is completely immersed in the water. This will prevent air bubbles being drawn into the microburette from the surface of the water.
3. Gently slide the plunger upwards, allowing about 0,2 to 0,3 ml of water to enter the microburette. Remove the container of water.
4. Slide the plunger of the syringe several times slowly up and down without forcing the water out of the microburette. The water will move up and down, rinsing the inside of the microburette. Discard the rinse solution into a large well of the comboplate®.
5. Repeat the rinsing step another two times.

B. Titrating with the microburette

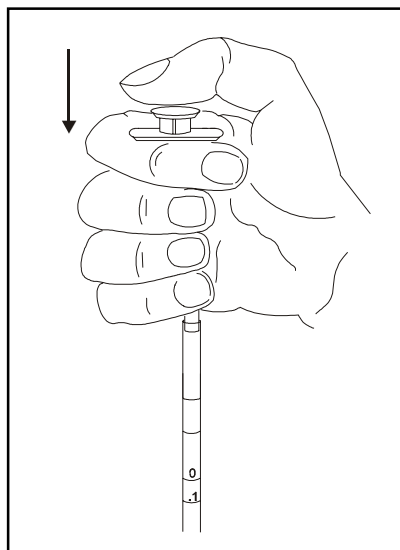
1. Immerse the end of the microburette as before into the water (or titrant solution). Gently slide the syringe plunger upwards to fill the burette up to the 1,00 ml level. **Do not pull the plunger up too quickly as this will cause the titrant solution to fill the entire microburette, including the silicone tubing and syringe.**
2. An air bubble may form at the end inside the plastic tip. This can be removed by forcing all of the titrant solution out of the microburette, and then refilling the microburette to the desired level. If the bubble persists, you may need to adjust or replace the plastic tip.
3. Read the bottom of the meniscus level and record the volume of the titrant solution in the microburette. Enter this as the initial volume in Table 1.

You may place a piece of white paper behind the microburette to make reading of the meniscus level easier.

TABLE 1

| Titration volume/ml | Titration 1 | Titration 2 | Titration 3 |
|---------------------|-------------|-------------|-------------|
| Final volume : | _____ | _____ | _____ |
| Initial volume : | _____ | _____ | _____ |
| Volume dispensed: | _____ | _____ | _____ |

4. To titrate, gently push the syringe plunger down until one drop of titrant has been dispensed. **Do not force the plunger down hard as this will cause a large volume of titrant to be dispensed. When performing an actual titration, this may lead to overshooting the end point.** Repeat this step until a total of five drops has been dispensed.



5. Record the volume of titrant remaining in the microburette by reading the bottom of the meniscus level. Enter this as the final volume in Table 1 and calculate the volume of titrant dispensed.
6. Practise the titration technique by repeating steps 1 - 5 twice more. Enter the results in Table 1.

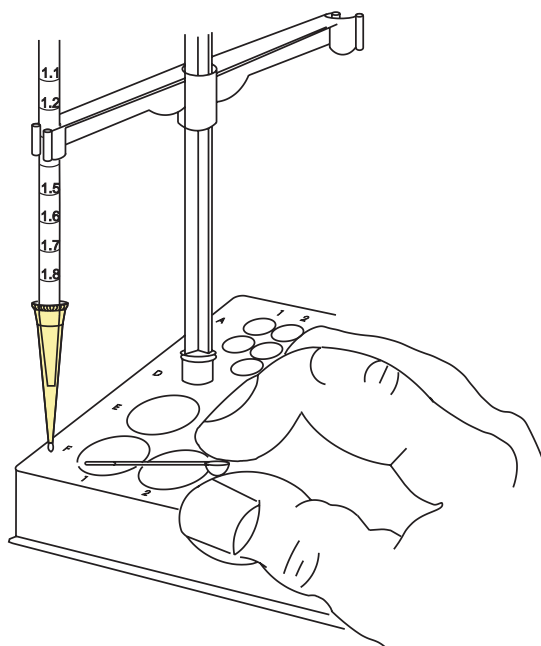
C. Dispensing fractional volumes using the microburette

When performing an actual titration, it may happen that the end point is reached before an entire drop of titrant is dispensed. In other words, less than one drop is needed to complete the titration. Fractional drops can be delivered using the technique below.

1. Push down very gently on the syringe plunger so that a drop of titrant begins to form at the end of the plastic tip, but is not released into the analyte solution.

This step may require special attention and practise as it involves fine control and manipulation of the syringe.

2. Touch the suspended droplet with the narrow end of the plastic microspatula. The droplet will adhere to the microspatula.



A CONDUCTOMETRIC TITRATION - THE DETERMINATION OF THE CONCENTRATION OF A ~ 0.01 M BARIUM HYDROXIDE SOLUTION, BY MEASURING THE CHANGE IN SOLUTION CONDUCTIVITY DURING A PRECIPITATION REACTION WITH 0.01 M SULPHURIC ACID

TEACHER'S GUIDE



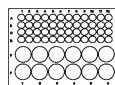
1. Chemicals

All the required chemicals are listed in the requirements table preceding the experiment. The teacher should obtain a standardised solution of sulphuric acid. If this is not possible, a solution of ~0,01 M $\text{H}_2\text{SO}_4(\text{aq})$ can be prepared by sufficiently diluting a more concentrated solution of the acid and standardising it against a standard sodium carbonate solution.

If 0.01 M $\text{Ba}(\text{OH})_2(\text{aq})$ is not available, it can be prepared by dissolving 0,7887g of solid barium hydroxide ($\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}(\text{s})$) in 250 ml boiled, cooled distilled water. The water must be allowed to cool to room temperature before preparing the $\text{Ba}(\text{OH})_2(\text{aq})$ because hot alkali solutions rapidly absorb carbon dioxide from the atmosphere, resulting in the formation of carbonates (in this case $\text{BaCO}_3(\text{s})$). The barium hydroxide solution should preferably be prepared on the same day that the titration is to be performed, due to the absorption of atmospheric carbon dioxide by the solution when it is in storage. The solid reagent from which the solution is prepared also absorbs $\text{CO}_2(\text{g})$, such that a solution is often contaminated with barium carbonate as soon as it has been prepared. If a white suspension or precipitate remains in the $\text{Ba}(\text{OH})_2(\text{aq})$ after making up the solution, then the solution must be filtered before use to remove the barium carbonate. It is also necessary to filter barium hydroxide solutions that have been in storage, especially when it is not possible to prepare fresh solutions each time the titration is performed. It should be remembered, however, that the longer the $\text{Ba}(\text{OH})_2(\text{aq})$ is stored, the more carbonate that will form and the more the concentration of the solution will be diminished.

The water of choice for preparation of the alkaline solution is also important. Distilled water is often supersaturated with carbon dioxide and must therefore be boiled and cooled to room temperature before dissolution of the solid occurs. Alternatively, deionised water can be used to prepare solutions because it contains negligible carbon dioxide.

If the $\text{Ba}(\text{OH})_2(\text{aq})$ is to be used over a prolonged period of time, it must be stored in a tightly stoppered polyethylene bottle and protected against the uptake of atmospheric carbon dioxide. Glass bottles are not recommended for long storage periods as barium hydroxide reacts slowly with the glass to form silicates. Stoppers for storage bottles containing $\text{Ba}(\text{OH})_2(\text{aq})$ should not be made of glass, because they will "freeze" upon prolonged exposure to the solution.



2. Equipment

Apart from the bar LED conductivity indicator, most of the apparatus required can be found in a RADMASTE Basic or Advanced Microchemistry Kit, and Microburette Kit. The paper cup (or any other suitable container) is used to catch the waste from the rinsing of the electrodes.

The bar LED conductivity indicator measures conductivity by displaying a series of lit bars (in the upper right corner of the comboplate®), when the electrodes of the instrument are in contact with a conductor or immersed in a conducting solution. It requires a 9V battery to function.



3. Hints

The teacher and students should follow or be familiar with the introductory notes on the construction and operation of the microburette prior to the experiment.

For this particular titration, the comboplate® of the bar LED must be used.

If the titration is performed on a traditional laboratory bench, the comboplate® may be raised to a more comfortable position using a sturdy, upturned box or by other suitable means. This prevents the student from having to bend down each time the microburette reading is taken. Alternatively, the student may sit down on a chair or low stool.

To deliver 1,00 ml aliquots of barium hydroxide solution into the wells of the comboplate®, copy the design of the microburette by attaching the plastic syringe to the top end of the other 2 ml pipette in the Microburette Kit with the silicone tubing, and place the plastic tip at the lower end of the pipette. Use this pipette to add exactly 1,00 ml of the $\text{Ba}(\text{OH})_2$ into the wells of the comboplate®.

The lid of the reagent bottle containing the $\text{Ba}(\text{OH})_2(\text{aq})$ must be replaced promptly after removing the 1,00 ml aliquot for analysis, otherwise the solution will be contaminated with carbonate as it absorbs atmospheric carbon dioxide.

The conductivity of the $\text{Ba}(\text{OH})_2(\text{aq})$ must be measured prior to adding any $\text{H}_2\text{SO}_4(\text{aq})$. This is done by immersing the electrode probe of the conductivity indicator in the centre of the $\text{Ba}(\text{OH})_2(\text{aq})$ in the well. The electrode wires at the end of the probe must be beneath the level of the solution in the well, but they must not be pushed against the bottom of the well. Such treatment will damage the electrodes.

It is best to wait a few seconds before recording the number of bars that are glowing. When the electrodes are immersed in the solution there may be some fluctuation in the conductivity of the solution, sometimes shown by one of the bars "flicking" on and off. The brief wait allows the conductivity indicator to settle so that a correct reading can be recorded.



The electrode probe must be promptly removed from the electrolyte solution, and any droplets of solution must be removed from the electrode wires as soon as the conductivity measurement has been made. This is done by touching the wires against the narrow edge of a microspatula so that the droplet adheres to the microspatula, which can then be used to transfer the droplet into the bulk of the solution. Similarly, when measuring the solution conductivity after acid addition to the $\text{Ba}(\text{OH})_2(\text{aq})$, it is extremely important to quickly remove the probe and remove any droplets of solution from the electrode wires. The bar LED runs on a Direct Current (DC) circuit. If the electrodes are allowed to remain in the electrolyte solution or if solution remains in contact with the electrodes after removing them from the well, they will become polarised and cause incoherent conductivity measurements to be recorded.

The 0,1 ml aliquots of $\text{H}_2\text{SO}_4(\text{aq})$ required may be added dropwise to maintain better control of titrant addition; for example, if 0,09 ml of acid has already been dispensed and the addition of another drop will result in more than 0,1 ml being added to the $\text{Ba}(\text{OH})_2(\text{aq})$, then a fraction of a drop of the $\text{H}_2\text{SO}_4(\text{aq})$ can be added using the end of the microspatula. As the titration proceeds and as more titrations are completed, this method of dropwise addition will become rapid.

If more than 0,1 ml $\text{H}_2\text{SO}_4(\text{aq})$ is added at any one time during the titration, the exact volume of acid dispensed must be recorded from the microburette. The exact volumes are required to draw conductivity curves for each titration.

After each 0,1 ml of $\text{H}_2\text{SO}_4(\text{aq})$ added, the solution in the well should be stirred. Stirring must not be vigorous, otherwise some of the solution may splash out of the well and introduce errors into the titration. The microspatula should be left in the well during the titration. This prevents loss of analyte solution and contamination, which may occur if the microspatula is removed.

When measuring the conductivity of the reaction solution after each aliquot of acid added, the electrode wires of the conductivity indicator must not be in contact with the microspatula in the well. The wires must be below the surface of the solution and a few seconds are needed until any fluctuation in conductivity has subsided.

For each titration, a graph of conductivity measured (number of bars lit) versus volume of $\text{H}_2\text{SO}_4(\text{aq})$ must be constructed using the data in Table 1. As the titration proceeds, the conductivity first decreases to a minimum. The first part of the graph will therefore have points that display a negative slope. After the end point of the titration, the conductivity increases from the minimum so that the latter part of the graph has points that display a positive slope. In order to find the true minimum point on the graph, the best straight line must be drawn through the points that show a negative slope (i.e. from 0,00 ml $\text{H}_2\text{SO}_4(\text{aq})$ to the volume of $\text{H}_2\text{SO}_4(\text{aq})$ at which the minimum conductivity was recorded). Another straight line must be drawn through the points that show a positive slope (i.e. from the volume of $\text{H}_2\text{SO}_4(\text{aq})$ at which the minimum conductivity was recorded to the volume of $\text{H}_2\text{SO}_4(\text{aq})$ at which the final conductivity was measured). The two lines will intersect at the true minimum conductivity of the titration solution.

A perpendicular must be extrapolated from the point of intersection onto the X axis to determine the volume of $\text{H}_2\text{SO}_4(\text{aq})$ required to reach minimum conductivity. From the stoichiometry of the reaction between $\text{H}_2\text{SO}_4(\text{aq})$ and $\text{Ba}(\text{OH})_2(\text{aq})$, the concentration of the $\text{Ba}(\text{OH})_2(\text{aq})$ can be calculated.

Students should be encouraged to observe what is happening inside the well as the titration proceeds, in order for the white barium sulphate precipitate to be identified.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

The ~0.01 M barium hydroxide solution (like all barium compounds) is poisonous if ingested. The solution must be handled with care. If any $\text{Ba}(\text{OH})_2(\text{aq})$ makes contact with the skin, wash the affected area with copious amounts of water. If possible, the mixture at the end of the experiment should be collected for disposal as heavy metal waste. If any base is squirted into the eye, immediately rinse the eye out under running water. Always have a dilute solution of boric acid close by to apply to the injury. This will help neutralise the base in the eye. The patient should be referred to a doctor.

Sulphuric acid is corrosive. If any acid is spilt on the skin, the affected area must immediately be rinsed with copious amounts of water. If any acid is squirted into the eye, immediately rinse the eye out under running water. Always have a dilute solution of sodium hydrogencarbonate (household baking soda), or milk close by to apply to the injury. These substances will help neutralise the acid in the eye. The patient should be referred to a doctor.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Example data are shown below.

Q1. Write down the exact $\text{H}_2\text{SO}_4(\text{aq})$ concentration which appears on the reagent bottle.

A1. **0.0111 M.**

Q2. Prepare a table like Table 1 below and record your observations in the table. Tabulate your result for titrations 1, 2 and 3.

A2. **TABLE 1: Change in solution conductivity as 1.00 ml $\text{Ba}(\text{OH})_2(\text{aq})$ is titrated with 0.01 M $\text{H}_2\text{SO}_4(\text{aq})$**

| titration number | 1 | | 2 | | 3 | |
|---|--|-------------------------------|--|-------------------------------|--|-------------------------------|
| Volume $\text{H}_2\text{SO}_4(\text{aq})$ /ml | Exact volume of $\text{H}_2\text{SO}_4(\text{aq})$ added /ml | Conductivity /no. of bars lit | Exact volume of $\text{H}_2\text{SO}_4(\text{aq})$ added /ml | Conductivity /no. of bars lit | Exact volume of $\text{H}_2\text{SO}_4(\text{aq})$ added /ml | Conductivity /no. of bars lit |
| 0.00 | 0.000 | 6 | 0.000 | 6 | 0.000 | 6 |
| 0.10 | 0.100 | 6 | 0.100 | 6 | 0.103 | 6 |
| 0.20 | 0.210 | 5 | 0.200 | 5 | 0.200 | 5 |
| 0.30 | 0.315 | 5 | 0.300 | 5 | 0.300 | 5 |
| 0.40 | 0.410 | 4 | 0.400 | 4 | 0.405 | 4 |
| 0.50 | 0.500 | 4 | 0.500 | 4 | 0.530 | 4 |
| 0.60 | 0.600 | 3 | 0.605 | 3 | 0.600 | 3 |
| 0.70 | 0.700 | 3 | 0.700 | 3 | 0.700 | 3 |
| 0.80 | 0.800 | 3 | 0.800 | 3 | 0.800 | 3 |
| 0.90 | 0.908 | 4 | 0.900 | 4 | 0.900 | 4 |
| 1.00 | 1.000 | 5 | 1.000 | 5 | 1.020 | 5 |
| 1.10 | 1.100 | 5 | 1.100 | 5 | 1.110 | 5 |
| 1.20 | 1.210 | 6 | 1.210 | 6 | 1.205 | 5 |
| 1.30 | 1.300 | 6 | 1.310 | 6 | 1.305 | 5 |
| 1.40 | 1.408 | 6 | 1.400 | 6 | 1.400 | 5 |
| 1.50 | 1.500 | 6 | 1.500 | 6 | 1.500 | 6 |
| 1.60 | - | - | - | - | - | - |

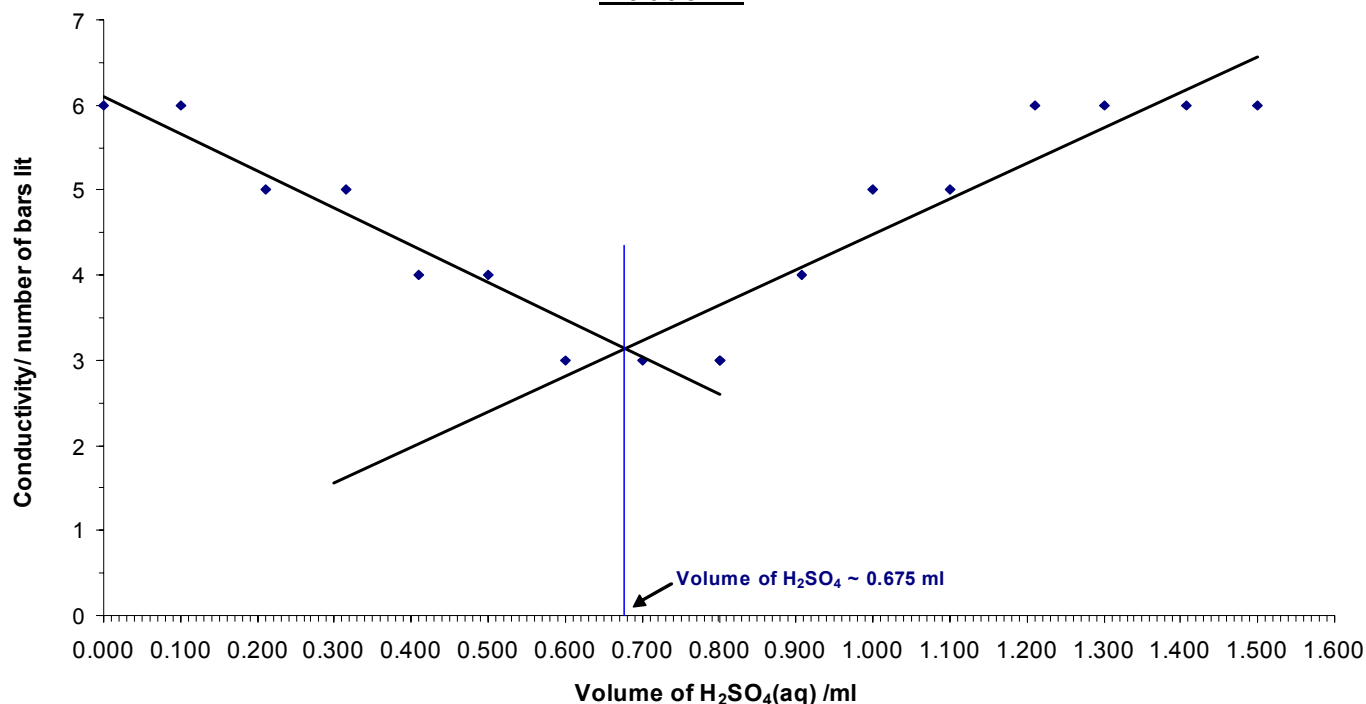
Q3. For each titration, prepare a graph of the conductivity/number of bars lit (Y axis) versus the volume of $\text{H}_2\text{SO}_4(\text{aq})$ (X axis). Plot the results obtained in Table 1 on the graph of each titration.

The scientific method used for finding the concentration of $\text{Ba}(\text{OH})_2(\text{aq})$ with graphs like the one which you have prepared, is to draw the best straight line through the set of points showing a negative slope and another best straight line through the set of points which display a positive slope. Therefore, draw the best straight line through the set of points between 0.00 ml and the volume of sulphuric acid at which the minimum number of bars was lit. Now draw the best straight line through the set of points between this volume that gave the lowest conductivity, and the volume of sulphuric acid at which you took your final measurement. Where the two lines intersect is the true minimum point on the curve (i.e. where the conductivity of the reaction solution was at its lowest.)

A3. **See Graph 1. (Note: only the graph for Titration 1 is shown. The graphs for Titrations 2 and 3 will be similar to Graph 1.)**

GRAPH 1: Graph of Conductivity vs Volume of H₂SO₄

Titration 1



Q4. Drop a perpendicular from this minimum point onto the X axis of each graph and record the volume of H₂SO₄(aq) where the perpendicular touches the axis.

A4. **Volume of H₂SO₄(aq) at minimum conductivity (titration 1) = 0.675 ml**

Volume of H₂SO₄(aq) at minimum conductivity (titration 2) = 0.675 ml

Volume of H₂SO₄(aq) at minimum conductivity (titration 3) = 0.665 ml

Q5. Write down the balanced chemical equation of the reaction between sulphuric acid and barium hydroxide.

A5. **H₂SO₄(aq) + Ba(OH)₂(aq) → BaSO₄(s) + 2H₂O(l)**

Q6. Use this equation and the volume of H₂SO₄(aq) at minimum conductivity obtained from each of the three graphs, to calculate the concentration of Ba(OH)₂(aq). Do not average the acid volume obtained from the three graphs; calculate the concentration of Ba(OH)₂(aq) from each set of titration data.

A6. **TITRATION 1:**

amount H₂SO₄(aq) = 0.0111 mol l⁻¹ x 0.675 x 10⁻³ l (Graph 1)

= 7.49 x 10⁻⁶ mol = moles Ba(OH)₂

Concentration of Ba(OH)₂

= $\frac{\text{amount Ba(OH)}_2}{\text{volume of Ba(OH)}_2 \text{ titrated}}$

= $\frac{7.49 \times 10^{-6} \text{ mol}}{1.00 \times 10^{-3} \text{ l}} = 7.49 \times 10^{-3} \text{ M}$

TITRATION 2:

The calculation for Titration 2 is the same as that for Titration 1 because the volume of H₂SO₄(aq) at minimum conductivity is also 0.675 ml.

TITRATION 3:

amount H₂SO₄(aq) = 0.0111 mol l⁻¹ x 0.665 x 10⁻³ l (Graph 3)

= 7.38 x 10⁻⁶ mol = moles Ba(OH)₂

Concentration of Ba(OH)₂

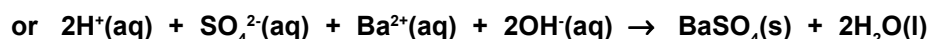
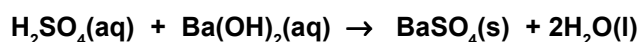
= $\frac{7.38 \times 10^{-6} \text{ mol}}{1.00 \times 10^{-3} \text{ l}} = 7.38 \times 10^{-3} \text{ M}$

- Q7. What change did you notice in the appearance of the reaction solution as sulphuric acid was added to the barium hydroxide?
- A7. The solution became cloudy and a white precipitate settled at the bottom of each well. As more sulphuric acid was added, the amount of precipitation increased.**
- Q8. Explain this change in appearance of the reaction solution. (Hint: look at the chemical equation you used in your calculations.)
- A8. As sulphuric acid ($\text{H}_2\text{SO}_4(\text{aq})$) reacted with the barium hydroxide ($\text{Ba}(\text{OH})_2(\text{aq})$) in the well, white barium sulphate ($\text{BaSO}_4(\text{s})$) formed. Barium sulphate is insoluble and therefore sank to the bottom of the well, causing the solution to appear cloudy as it precipitated.**
- Q9. What changes in the conductivity (number of bars lit) of the reaction solution did you observe whilst performing each titration?
- A9. At the start of each titration the conductivity of the $\text{Ba}(\text{OH})_2(\text{aq})$ in the well was high, with 6 bars lighting up when the electrode probe was immersed in the solution. As $\text{H}_2\text{SO}_4(\text{aq})$ was added the conductivity of the reaction solution began to decrease, as shown by a reduction in the number of bars lighting up. Eventually a minimum of 3 bars was lit, after which the conductivity began to increase as more aliquots of $\text{H}_2\text{SO}_4(\text{aq})$ were added.**

The titration was terminated when the conductivity had increased to a constant value of 6 glowing bars.

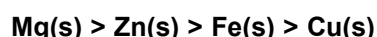
NOTE: the number of bars glowing at the beginning and end of the titration, as well as the number of bars glowing at minimal solution conductivity may differ from those stated above. Students must be able to deduce from their observations that the conductivity of the solution first decreases, reaches a minimum and then increases again.

- Q10. Explain the conductivity changes in terms of the reaction between $\text{H}_2\text{SO}_4(\text{aq})$ and $\text{Ba}(\text{OH})_2(\text{aq})$.
- A10. Prior to adding $\text{H}_2\text{SO}_4(\text{aq})$, the well contains only $\text{Ba}(\text{OH})_2(\text{aq})$. The solution therefore contains $\text{Ba}^{2+}(\text{aq})$ and $\text{OH}^-(\text{aq})$ ions, and conducts a current causing a high number of bars to light up. As $\text{H}_2\text{SO}_4(\text{aq})$ is added to the $\text{Ba}(\text{OH})_2(\text{aq})$, the following reaction occurs:**



Some of the Ba^{2+} ions bond with the SO_4^{2-} ions of the sulphuric acid to form insoluble barium sulphate. Some of the OH^- ions bond with the H^+ ions of the acid to form water, which is also non-conducting. Since some of the conducting ions have been removed from the solution, the conductivity decreases and less of the bars glow. As more $\text{H}_2\text{SO}_4(\text{aq})$ is added, more $\text{BaSO}_4(\text{s})$ and $\text{H}_2\text{O}(\text{l})$ are formed until all (or almost all) of the conducting ions in the solution are depleted. The solution conductivity reaches a minimum where the least number of bars light up. Beyond the end point, there are no more $\text{Ba}^{2+}(\text{aq})$ and $\text{OH}^-(\text{aq})$ ions remaining in the solution. As more $\text{H}_2\text{SO}_4(\text{aq})$ is added, the well contains only $\text{H}^+(\text{aq})$ ions, $\text{SO}_4^{2-}(\text{aq})$ ions, solid barium sulphate and water. Because of the presence of these ions, the solution conducts and the number of glowing bars begins to increase. The greater the volume of acid added after the end point, the greater the concentration of conducting ions present in the solution and the higher the conductivity of the solution.

- Q11. If the conductivity indicator was extremely sensitive, what conductivity would you expect to measure at the end point of the titration?
- A11. The expected solution conductivity would be zero.**
- Q12. How would you detect this conductivity?
- A12. None of the bars on the conductivity indicator would glow.**
- Q13. Explain why you would get this conductivity.
- A13. At the end point, the stoichiometric volumes of $\text{H}_2\text{SO}_4(\text{aq})$ and $\text{Ba}(\text{OH})_2(\text{aq})$ react to give only $\text{BaSO}_4(\text{s})$ and $\text{H}_2\text{O}(\text{l})$. There are no conducting ions remaining in the solution (except the very low (10^{-7} M) concentration of H^+ and OH^- ions that occur in pure water) and therefore zero conductivity.**
- Q14. What metals/non-metals are suitable for electrodes that measure conductivity? Which metals are unsuitable for electrodes?
- A14. Copper is a suitable metal. Metals such as magnesium, iron and zinc are good conductors, but not usually suitable for electrodes because of their high reactivity as depicted by the metal reactivity series:**



Platinum is an excellent electrode material because it is an inert metal. Other metals may also be platinised (i.e covered with a thin layer of platinum) to be used as electrodes. Carbon (graphite), although a non-metal, is a good conductor and fairly inert. It is therefore often used for electrodes.

- Q15. What is the effect of the distance between the electrodes on the number of lit bars?
- A15. The closer together the electrodes are, the lower the resistance in the circuit. Since conductivity is the reciprocal of resistance, the conductivity increases as the distance between the electrodes decreases. This means that more bars will light up the closer together the electrodes are.**

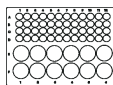
POTENTIOMETRIC TITRATION OF AN ACID AND A BASE

TEACHER'S GUIDE



1. Chemicals

All of the required chemicals are listed in the worksheet.



2. Equipment

Most of the apparatus required can be found in a RADMASTE Advanced Microchemistry Kit and Microburette Kit. A multimeter (or voltmeter) and connecting wires for the multimeter are required. For this experiment it is essential that the multimeter used is accurate. The Sellotape is not supplied in the kit.



3. Hints

It is not necessary for all learners to prepare the acidified potassium permanganate solution to activate the carbon/graphite rod. 3 or 4 carbon/graphite rods can be activated in one well. The glass rod used to hold the carbon rods in place allows for a constant distance between the rods during titrations, and also serves as a support for clamping the electrode assembly. The pipette must be thoroughly rinsed to obtain accurate results. Please ask the learners to follow the notes carefully when assembling the microburette.

It must be explained to the learner that if the wires from the multimeter are not connected correctly the reading on the multimeter will be negative. If the reading on the multimeter is negative, disconnect the copper wires from the multimeter and attach the wires to the multimeter again so that the reading on the multimeter is positive. If the multimeter records zero, then the connection of all wires is not done properly. Make sure all wires are connected.

Before commencing with the titration, make sure the two graphite rods placed in the well containing the HCl solution are at the same level. Please remind the learners to wash all used equipment immediately after the experiment has been completed.



4. Cautions

Please remember the following cautions and inform your students of all safety hazards:

Make sure that all learners wash their hands thoroughly after the experiment, as the chemicals can be irritating to the skin.

Never point a pipette or a syringe containing acid or base upwards. A momentary lapse of concentration can result in a nasty accident. If any acid or base is squirted into the eye, immediately rinse the eye out under running water. In the case of an acid, always have a dilute solution of sodium hydrogencarbonate (household baking soda), or milk close by to apply to the injury. In the case of a base, apply a dilute boric acid solution to the injury. These substances will help neutralise the acid or base in the eye. The patient should be referred to a doctor.



5. Model Answers to Questions in the Worksheet

It is recommended that learners write down all of the questions and answers in their workbooks. If this is done, then the answers to questions do not have to be in full sentences. If the learners do not copy the questions into their workbooks, then answers should be written in full sentences. Note that some of the questions can only be answered by learners in higher grades. Word equations can be written instead of chemical equations where required.

Q1. What is the reading (mV) on the multimeter? Construct a Table like the one below, and enter the multimeter reading and the volume of NaOH(aq).

A1. **The reading will vary, depending on how well the carbon/graphite rod is activated. The reading can range from 100 mV to 700 mV. What is of importance in this experiment is for learners to see when the approximate end-point is reached (Titration 1) and then graphically determine the end-point (Question 3).**

TABLE 1: Change in multimeter reading (mV) as $\text{HCl}(\text{aq})$ is titrated with $\text{NaOH}(\text{aq})$

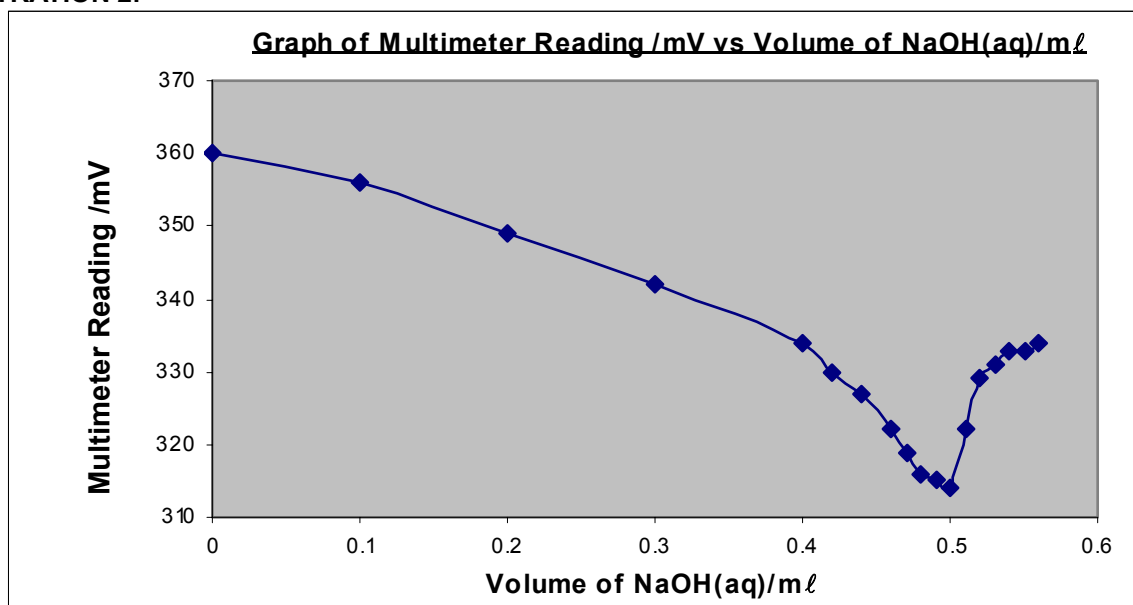
| Volume $\text{NaOH}(\text{aq})$ / $\text{m}\ell$ | MULTIMETER READING (mV) | | |
|--|-------------------------|-------------|-------------|
| | TITRATION 1 | TITRATION 2 | TITRATION 3 |
| 0.000 | 430 | 360 | 122 |
| 0.100 | 420 | 356 | 120 |
| 0.200 | 420 | 349 | 116 |
| 0.300 | 440 | 342 | 105 |
| 0.400 | 450 | 334 | 92 |
| 0.420 | | 330 | 85 |
| 0.440 | | 327 | 79 |
| 0.460 | | 322 | 75 |
| 0.470 | | 319 | 73 |
| 0.480 | | 316 | 70 |
| 0.490 | | 315 | 68 |
| 0.500 | 420 | 314 | 66 |
| 0.510 | | 322 | 75 |
| 0.520 | | 329 | 83 |
| 0.530 | | 331 | 87 |
| 0.540 | | 333 | 90 |
| 0.550 | | 333 | 90 |
| 0.560 | | 334 | 91 |

Q2. In Titration 1, at what volume of $\text{NaOH}(\text{aq})$ was the approximate end-point reached according to the colour change of the universal indicator?

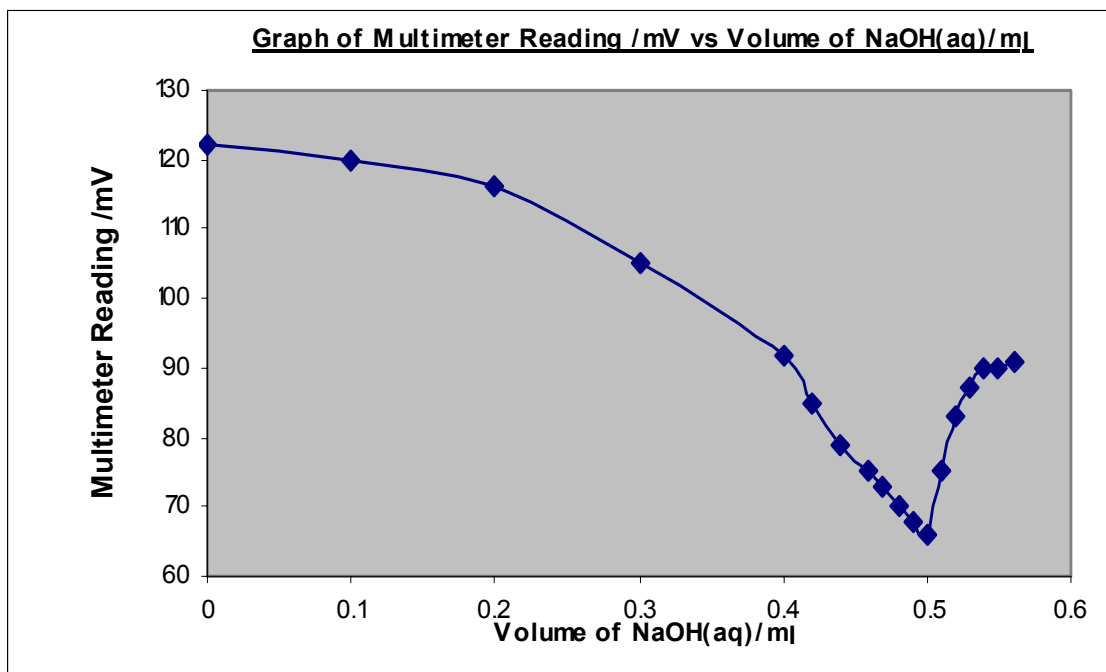
A2. At approximately 0.50 $\text{m}\ell$ of $\text{NaOH}(\text{aq})$.

Q3. Draw a graph of multimeter reading (mV) vs volume of $\text{NaOH}(\text{aq})$ ($\text{m}\ell$) by using the readings for either Titration 2 or 3.

A3. TITRATION 2:



TITRATION 3:



Q4. What is the volume of NaOH(aq) at the end point of the titration ?

A4. The volume at the end-point of the titration is 0.500 ml.

Q5. If the hydrochloric acid used was exactly 0.1000 M , what would be the concentration of the NaOH(aq) ?

A5. The end-point occurred when 0.500 ml of NaOH(aq) was titrated with exactly 0.500 ml of HCl(aq). Therefore if the concentration of HCl(aq) used was 0.1000 M, then the concentration of NaOH(aq) would be 0.1000 M exactly.

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