Cambridge Assessment International Education

Teaching Pack

Constructing electrochemical cells and measuring electrode potentials

Cambridge International AS & A Level Chemistry 9701

Version 1.0

Cambridge Pathway

Contents

Icons used in this pack:

Briefing lesson

Planning lesson

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Introduction

This pack will help you to develop your learners' experimental skills as defined by assessment objective 3 (AO3 Experimental skills and investigations) in the course syllabus.

Important note
Our *Teaching Packs* have been written by **classroom teachers** to help you deliver topics and skills that can be challenging. Use these materials to supplement your teaching and engage your learners. You can also use them to help you create lesson plans for other experiments.

This content is designed to give you and your learners the chance to explore practical skills. It is not intended as specific practice for Paper 3 (Advanced Practical Skills) or Paper 5 (Planning, Analysis and Evaluation).

This is one of a range of *Teaching Packs* and each pack is based on one experiment. The packs can be used in any order to suit your teaching sequence.

The structure is as follows:

Briefing lesson (1 hour*)
This lesson reinforces the key concepts, relevant skills, knowledge and understanding linked to the experiment.

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Planning lesson (1 hour*)
This lesson focuses on planning an experiment.
It ends with a demonstration video of the experiment.

↓

Lab lesson (1 hour*)
This lesson focuses on carrying out the experiment including the collection and recording of observations, measurements and estimates.

↓

Debriefing lesson (1 hour*)
This lesson focuses on the analysis and interpretation of data.
This includes making conclusions, evaluating methods and the quality of data and how improvements could be made.

** the timings are a guide only; you may need to adapt the lessons to suit your circumstances.*

In this pack you will find lesson plans, worksheets and teacher resource sheets.

Experiment: Constructing an electrochemical cell and measuring electrode potentials

This *Teaching Pack* focuses on an electrochemical cell experiment

Electrochemical cells are widely used. During this experiment, learners will construct an electrochemical cell and predict the cell potential using given data.

This experiment has links to the following syllabus content (see syllabus for detail):

- 24.2 Standard electrode potentials

The experiment covers the following experimental skills, as listed in **AO3: Experimental skills and investigations:**

- plan experiments and investigations
- collect, record and present observations, measurements and estimates
- analyse and interpret data to reach conclusions
- evaluate methods and quality of data and suggest improvements.

Prior knowledge

Knowledge from the following syllabus topics is useful for this experiment.

- 6.1 Redox processes
- 24.1 Electrolysis

Briefing lesson: Electrochemical Cells

Resources	• Image of an electrochemical cell • Worksheets A, B, C and D

Learning objectives	By the end of the lesson: • ***all*** learners should be able to describe an electrochemical cell • ***most*** learners should be able to state what *electrode potential* means • ***some*** learners will be able to predict reactions at electrodes based on electrode potential.

Timings	Activity
10 min	**Starter/Introduction** Ask learners to complete the redox equations from Worksheet A.
10 min	**Main lesson** Ask learners to discuss how the redox equations from the starter can be used. Lead them to chemical cells and batteries. How many reactions would be needed in the battery? Why are they called half-equations? Show an image of an electrochemical cell – what do learners think will happen? Hand out Worksheet B. Ask learners to describe what is happening. Do they notice any patterns between oxidation/reduction and the reactivity of metals? Is one metal always oxidised? How are the cell formulae written? How is this different from the half-equation? If learners are struggling – lead them to reactivity.
20 min	Lead the lesson on to electrode potentials. Ask learners what does the term *electrode potential* sound like? Lead to the idea that it is the potential of the electrode to lose/gain electrons under standard conditions when connected to a hydrogen half-cell. The more positive the number the better at gaining electrons (will be reduced), the more negative the better at losing electrons (will be oxidised). What would happen if an electrochemical cell had two half-cells with the same electrode potential? Hand out Worksheet C – learners decide, based on electrode potentials, which substance would be oxidised and which would be reduced and write half-equations.
10 min	Go through the answers – lead to the image at the end of Worksheet C – what is the link between the potential difference of the cell and the two electrode potentials. Ask learners to think how those numbers are related. Lead to the equation to calculate E_{cell} (more positive electrode potential minus less positive electrode potential). Learners complete Worksheet D.
10 min	**Plenary** Ask learners to draw and label a diagram of an electrochemical cell that would produce the biggest potential difference using the electrode potentials from earlier in the lesson and a cell that would produce the smallest potential difference, bigger than zero.

Planning lesson: Investigating electrochemical cells

Resources	• List of standard electrode potentials as examples • Experiment video • Worksheet E

Learning objectives	By the end of the lesson: • ***all*** learners should be able to plan an electrochemical cell experiment • ***most*** learners should be able to describe the limitations of the electrode potential numbers • ***some*** learners will be able to analyse experimental methods.

<table>
<tr><th>Timings</th><th>Activity</th></tr>
<tr><td>10 min</td><td>Starter/Introduction
Ask learners how they would set up an electrochemical cell. What safety precautions would they need to consider? If they were to use it to prove the electrode potential values, what other conditions would they need to consider? (standard conditions)</td></tr>
<tr><td>25 min

5 min

10 min</td><td>Main lesson
Learners plan an experiment to prove that the electrode potentials of different half-cells effect which substance is oxidised and which is reduced. (This will involve creating different electrochemical cells with different combinations of half-cells and finding which is the positive electrode (will be reduced) and which is the negative electrode (will be oxidised). Displaying a list of standard electrode potentials would give the learners ideas of which metals to use. They need to consider:
• what equipment is needed
• examples of metals and salt solutions they could use
• any safety rules they may need to follow
• how would their results prove that electrode potentials of the two half-cells make a difference?
• how could they tell which substance is oxidised by observations?
• how could they tell which substance is oxidised by potential difference recorded (negative electrode oxidised when V is a positive number)
• would they need to follow standard conditions?

Learners should check each other’s work to see if they have missed anything or if anything does not make sense.

Watch the Constructing an electrochemical cell and measuring electrode potentials experiment video. Discuss with learners if they forgot to mention anything in their plan. Ask learners if they can think of any problems with the electrode potentials. Lead to the idea they are under standard conditions. How often does this happen? What might affect the electrode potential? Lead to ideas of non-standard conditions, e.g. concentration. Learners plan an experiment to test the effect of non-standard conditions (e.g. different concentrations).</td></tr>
<tr><td>10 min</td><td>Plenary
Worksheet E – learners analyse an experimental plan. What is wrong/right with it?</td></tr>
</table>

Lab lesson: Electrochemical cell experiment

Resources

- *Teacher walkthrough* video
- Experimental equipment
- Teacher notes
- Worksheet F (if required)

Learning objectives

By the end of the lesson:

- ***all*** learners should be able to set up an electrochemical cell
- ***most*** learners should be able to collect results for an electrochemical cell experiment
- ***some*** learners will be able to predict the pattern of results based on electrode potentials.

Timings	Activity
10 min	**Starter/Introduction** Ask learners to explain how to set up an electrochemical cell – you could do this as a demo where they tell you what to do with the given equipment or you could watch the experiment video again with no sound and ask them to explain what is happening.
40 min	**Main lesson** Learners should design a table for their results. Learners carry out the experiment based on their plan from last lesson. They should also make a note of which electrode is the positive one each time. **Safety** Circulate the classroom at all times during the experiment so that you can make sure that your learners are safe and that the data they are collecting is accurate. If learners need a bit more guidance give them the learner method on Worksheet F. Learners should use different half-cells to prove the electrode potential effects potential difference. Can they predict if the potential difference will be higher or lower based on which half-cells they are using? What happens to the potential difference over time?
10 min	**Plenary** Learners should write a conclusion about their results. Were there any anomalies? Did they get what they expected? Were the potential differences the same number as the calculated values – why would this be? What happens to the potential difference if the cell is left on? What can we relate this to in real life? (batteries)

Teacher notes

Watch the electrochemical cell video (*Teacher Walkthrough* version) and read these notes.

Each group will require:
- access to saturated sodium chloride solution
- metal rods and metal salt solutions
- two 100 cm^3 beakers
- strips of filter paper
- two crocodile clips
- two wires
- a voltmeter.

Metals and solutions you could use include: zinc and zinc sulfate, copper and copper sulfate, magnesium and magnesium sulfate, iron and iron sulfate.

Safety

The information in the table below is a summary of the key points you should consider before undertaking this experiment with your learners.

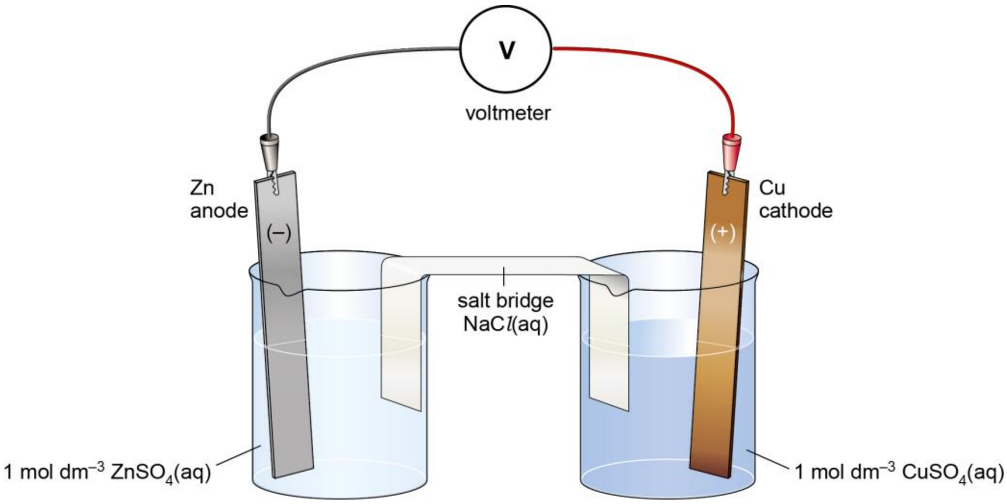
It is your responsibility to carry out an appropriate risk assessment for this experiment.

Substance	Hazard	First aid
Sodium chloride [saturated solution]	Low hazard	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor. **If inhaled:** move person into fresh air. If not breathing, give artificial respiration. Consult a doctor. **Spilt on skin or clothing:** remove contaminated clothing and shoes immediately and rinse. Wash off the skin with plenty of water. Consult a doctor.
Copper sulfate [1 mol dm^{-3} solution]	 GHS07 (*moderate hazard* **MH**)  GHS09 (*hazardous to the aquatic environment* **N**)	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor. **If inhaled:** move person into fresh air. If not breathing, give artificial respiration. Consult a doctor. **Spilt on skin or clothing:** remove contaminated clothing and shoes immediately and rinse. Wash off the skin with plenty of water. Consult a doctor.

Substance	Hazard	First aid
Copper [metal]	Low hazard	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor.
Magnesium sulfate [1 mol dm^{-3} solution]	Low hazard	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor. **If inhaled:** move person into fresh air. If not breathing, give artificial respiration. Consult a doctor. **Spilt on skin or clothing:** remove contaminated clothing and shoes immediately and rinse. Wash off skin with plenty of water. Consult a doctor.
Magnesium [metal]	 GHS02 (flammable F)	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor.
Zinc sulfate [1 mol dm^{-3}solution]	 GHS05 (*corrosive* **C**)  GHS07 (*moderate hazard* **MH**)  GHS09 (*hazardous to the aquatic environment* **N**)	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor. **If inhaled:** move person into fresh air. If not breathing, give artificial respiration. Consult a doctor. **Spilt on skin or clothing:** remove contaminated clothing and shoes immediately and rinse. Wash off the skin with plenty of water. Consult a doctor.
Zinc [metal]	Low hazard	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor.

Substance	Hazard	First aid
Iron sulfate [1 mol dm^{-3} solution]	 GHS07 (*moderate hazard* **MH**)	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor. **If inhaled:** move person into fresh air. If not breathing, give artificial respiration. Consult a doctor. **Spilt on skin or clothing:** remove contaminated clothing and shoes immediately and rinse. Wash off the skin with plenty of water. Consult a doctor.
Iron [metal]	Low hazard	**In the eye:** rinse thoroughly with plenty of water for at least 15 min and consult a doctor. **Swallowed:** wash out the mouth with water. Do not induce vomiting. Never give anything by mouth to an unconscious person. Consult a doctor.

Experiment set-up



Teacher method

This is your version of the method for this experiment that accompanies the *Teacher walkthrough* video. Do not share this method with learners.

Before you begin

Plan how you will group your learners during the experiment session.

Think about:

- the number of groups you will need (group size 2–4 learners)
- the amount of equipment/chemicals required
- whether you are testing more than one electrochemical cell or not

Experiment

Walk around the learners during the experiment in case they encounter any difficulties.

Steps	Notes
1. Learners should collect the equipment they require from the front of the class.	*Ensure students know how to use a voltmeter, how to swap wires to get a positive reading and that this shows which electrode is the positive electrode.*
2. They should find a space in the classroom where the equipment can be assembled safely.	
3. Make sure your learners know how to set up the experiment.	*Learners may need help with the salt bridge.*
4. Learners should create two different half-cells of metal rods in the same metal salt solutions.	
5. Learners should create a salt bridge by soaking the filter paper in the saturated $NaCl$ and putting the paper so it connects the two half-cells like a bridge.	
6. Learners record the potential difference of the cell.	
7. Learners repeat with different half-cells.	

Clean-up

After the experiment learners should:

- clean all glassware
- tidy up their work space
- ensure any spillages have been mopped up
- return all equipment and any unused chemicals to you.

Alternative methods

If you do not have access to the required equipment or the suggested method would not work for your class, here are some possible alternatives that you could use.

You could perform the experiment with just one electrochemical cell if you don’t have multiple examples of chemicals and metals. You could also investigate the potential difference change over time or measure the potential difference with the same chemical at different concentrations.

Debriefing lesson: Electrochemical cell analysis

Resources	• Worksheet G • Worksheet H • List of standard electrode potentials

Learning objectives	By the end of the lesson: • ***all*** learners should be able to analyse the method used to make electrochemical cells • ***most*** learners should be able to analyse their results from the practical • ***some*** learners will be able to answer questions based on electrochemical cell experiments.

Timings	**Activity**
10 min	**Starter/Introduction** Learners write a list of what made the electrochemical cell practical easy and what problems there were with it. Discuss as a class and if any problems with the experiment do come up, can other learners solve them?
20 min 20 min	**Main lesson** Give learners Worksheet G. Learners use the prompts from the worksheet and look at their results from the experiment; do their results fit the pattern? Did any cell not work? Why might this have been? Did any cell give an unexpected reading? Learners should make a list of the lowest electrode potential to the highest using their data. (The positive electrode from their results will have the more positive electrode potential in each of the cells the learners made). Does their list fit with the given list at standard conditions? Why/why not? e.g. did they use different concentrations? Give learners Worksheet H – questions on electrochemical cells. When learners are finished, get them to go through the answers with each other – provide help where required.
10 min	**Plenary** Learners write a method for an electrochemical cell experiment with some deliberate mistakes. Ask a fellow learner to correct it.

Worksheets and answers

Worksheet A: Redox half-equations

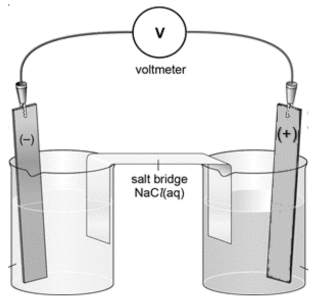
In your lab book, complete the following half-equations:

1. $Mg \rightarrow Mg^{2+}$
2. $Fe \rightarrow Fe^{2+}$
3. $Cl^{-} \rightarrow Cl_2$
4. $Cu^{2+} \rightarrow Cu$
5. $Ag^{+} \rightarrow Ag$
6. $Fe \rightarrow Fe^{3+}$
7. $O^{2-} \rightarrow O_2$
8. $Al \rightarrow Al^{3+}$
9. $Fe^{2+} \rightarrow Fe$
10. $Zn \rightarrow Zn^{2+}$
11. Combine the half-equations from questions 5 and 6
12. Combine the half-equations from questions 4 and 8

Worksheet B: Electrochemical cells

Look at the information below regarding different electrochemical cells.

1. Label each half-equation as either *oxidation* or *reduction*.
2. What is the difference between the half-equation and the way the cell is written as a half-cell?
3. Can you see any link between the metals in the half-cells and which is oxidised?
4. Can you see any link between the metals in the half-cell and which is the positive electrode?



Zinc–copper electrochemical cell				
Electrodes	**Charge of electrode**	**Half-equation**	**Oxidation or reduction?**	**Half-cell**
zinc	−	$Zn \rightarrow Zn^{2+} + 2e^-$		Zn^{2+} (aq) | Zn (s)
copper	+	$Cu^{2+} + 2e^- \rightarrow Cu$		Cu^{2+} (aq) | Cu (s)

Magnesium–copper electrochemical cell				
Electrodes	**Charge of electrodes**	**Half-equation**	**Oxidation or reduction?**	**Half-cell**
magnesium	−	$Mg \rightarrow Mg^{2+} + 2e^-$		Mg^{2+} (aq) | Mg (s)
copper	+	$Cu^{2+} + 2e^- \rightarrow Cu$		Cu^{2+} (aq) | Cu (s)

Silver–copper electrochemical cell				
Electrodes	**Charge of electrode**	**Half-equation**	**Oxidation or reduction**	**Half-cell**
silver	+	$Ag^+ + e^- \rightarrow Ag$		Ag^+ (aq) | Ag (s)
copper	−	$Cu \rightarrow Cu^{2+} + 2e^-$		Cu^{2+} (aq) | Cu (s)

Worksheet C: Electrode potentials

Electrode potential ($E^{\ominus}$) is the potential of the electrode to lose or gain electrons under standard conditions. The more positive the value the better at gaining electrons the metal/element will be (and so the element will be reduced), the more negative the value, the better at losing electrons the metal will be (and so the metal will be oxidised). **Remember** the negative electrode is oxidised.

Look at the following electrochemical cells.

1. Use the electrode potentials to decide which metal is oxidised and which is reduced.
2. Use this to also decide which is the positive and negative electrode.
3. Write half-equations for the oxidation and reduction reactions.

Half-cell	$E^{\ominus}$ (V)
Ag^{+}(aq) | Ag(s)	+0.80
Fe^{2+}(aq) | Fe(s)	−0.44

Half-cell	$E^{\ominus}$ (V)
Cl_2(g) | Cl^{-}(aq)	+1.36
I_2(g) | $2I^{-}$(aq)	+0.54

Half-cell	$E^{\ominus}$ (V)
Mg^{2+}(aq) | Mg(s)	−2.37
Au^{+}(aq) | Au(s)	+1.68

Half-cell	$E^{\ominus}$ (V)
Cl_2(g) | Cl^{-}(aq)	+1.36
Au^{+}(aq) | Au(s)	+1.68

Half-cell	$E^{\ominus}$ (V)
Zn^{2+}(aq) | Zn(s)	+0.76
Al^{3+}(aq) | Al(s)	−1.66

Half-cell	$E^{\ominus}$ (V)
Cl_2(g) | Cl^{-}(aq)	+1.36
$2H^{+}$(aq) | H_2(g)	0.00

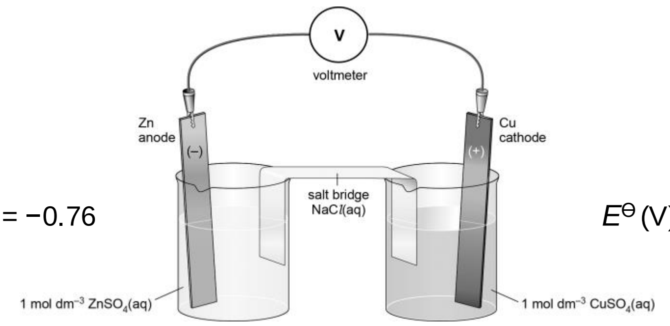
Half-cell	$E^{\ominus}$ (V)
Au^{+}(aq) | Au(s)	+1.68
$2H^{+}$(aq) | H_2(g)	0.00

Half-cell	$E^{\ominus}$ (V)
Mg^{2+}(aq) | Mg(s)	−2.37
I_2(g) | $2I^{-}$(aq)	+0.54

What do you notice about the following image?

Reading from voltmeter = +1.10 V

$E^{\ominus}$ (V) = −0.76



$E^{\ominus}$ (V) = +0.34

Worksheet D: E_{cell}

E_{cell} = more positive electrode potential minus less positive electrode potential.

Use the table to help you complete the following questions.

Half-equation	$E^{\ominus}$ / V
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$	−2.37
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	−1.66
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	−0.76
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	−0.44
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	−0.14
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$	+0.15
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34
$Cu^+(aq) + e^- \rightarrow Cu(s)$	+0.52
$I_2(aq) + 2e^- \rightarrow 2I^-(aq)$	+0.54
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$Ag^{2+}(aq) + e^- \rightarrow Ag^+(aq)$	+1.98

1. Calculate E_{cell} for the following electrochemical cells:

   a. $Cu^{2+}(aq) \mid Cu(s)$ and $Mg^{2+}(aq) \mid Mg(s)$
   b. $Ag^+(aq) \mid Ag(s)$ and $Mg^{2+}(aq) \mid Mg(s)$
   c. $Fe^{2+}(aq) \mid Fe(s)$ and $Al^{3+}(aq) \mid Al(s)$
   d. $Pb^{2+}(aq) \mid Pb(s)$ and $Cu^+(aq) \mid Cu(s)$
   e. $Zn^{2+}(aq) \mid Zn(s)$ and $Pb^{2+}(aq) \mid Pb(s)$
   f. $Cu^{2+}(aq) \mid Cu(s)$ and $Br_2(g) \mid 2Br^-(aq)$
   g. $Fe^{2+}(aq) \mid Fe(s)$ and $Ag^{2+}(aq) \mid Ag^+(aq)$
   h. $Al^{3+}(aq) \mid Al(s)$ and $Br_2(g) \mid 2Br^-(aq)$
   i. $Cu^{2+}(aq) \mid Cu(s)$ and $Cu^+(aq) \mid Cu(s)$
   j. $Pb^{2+}(aq) \mid Pb(s)$ and $Mg^{2+}(aq) \mid Mg(s)$

2. Using the data above decide whether these reactions will occur under standard conditions. Remember to look at which substance is being reduced in the equation; if $E^{\ominus}$ is more positive, the substance should be reduced.

   a. $Ag^+(aq) + Fe^{2+}(aq) \rightarrow Ag(s) + Fe^{3+}(aq)$
   b. $Cu(s) + Pb^{2+}(aq) \rightarrow Cu^{2+}(aq) + Pb(s)$
   c. $2Fe^{3+}(aq) + 2Br^-(aq) \rightarrow 2Fe^{2+}(aq) + Br_2(aq)$
   d. $2Cu^+(aq) \rightarrow Cu^{2+}(aq) + Cu(s)$

3. Use the data above to predict whether the following substances will react together under standard conditions and give the full equation if a reaction is expected.

   a. Pb with $H^+(aq)$
   b. Cu with $H^+(aq)$

Worksheet E: Experimental plan to analyse

Constructing an Electrochemical Cell

Read through this method and see if you can find any problems. Think about whether you could follow this method.

Equipment needed

- saturated sodium chloride solution
- metal rods and metal salt solutions (zinc and zinc sulfate, copper and copper sulfate, magnesium and magnesium sulfate, iron and iron sulfate)
- two 100 cm^3 beakers
- strips of filter paper
- a voltmeter.

Method

- Pour 50 cm^3 of salt solution into a beaker. Add a metal rod.
- Pour 50 cm^3 of another salt solution into another beaker. Add a metal rod.
- Connect a crocodile clips to your metal electrodes and connect these to the voltmeter.
- Use the filter paper to create a salt bridge between the two half-cells.
- Record the potential difference.
- Repeat with different combinations of half-cells.

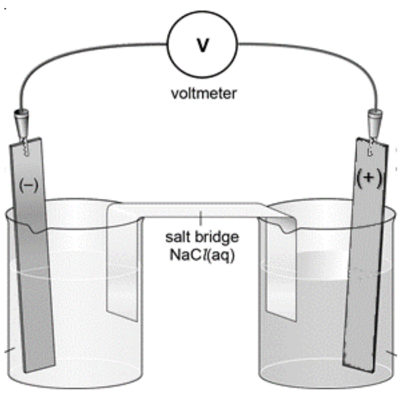
Safety

Some materials are flammable.

Errors

The reading will change over time as the reaction continues, so if the reading is not taken at a similar time for each reaction this could lead to different results.

Diagram of experimental setup



Worksheet F: Learner method

Equipment

- saturated sodium chloride solution
- metal rods and 1 M metal salt solutions (zinc and zinc sulfate, copper and copper sulfate, magnesium and magnesium sulfate, iron and iron sulfate)
- two 100 cm^3 beakers
- strips of filter paper
- a voltmeter
- two crocodile clips
- two wires.

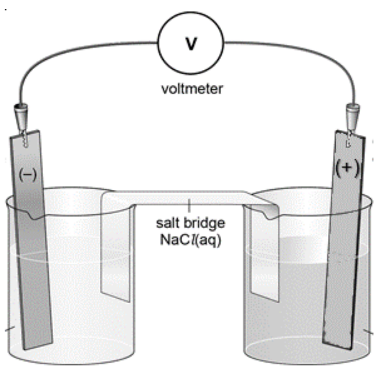
Method

- Pour 50 cm^3 of any salt solution into one of the beakers. Place the corresponding metal into the salt solution (e.g. zinc and zinc sulfate). This is one of your half-cells.
- Pour 50 cm^3 of another salt solution into the other beaker. Place the corresponding metal into the salt solution (e.g. iron and iron sulfate). This is the other half-cell.
- Connect a crocodile clip to one of your metal electrodes and connect this via a wire to the voltmeter.
- Connect a crocodile clip to your other metal electrode and connect this via a wire to the voltmeter.
- Soak the strip of filter paper in the saturated sodium chloride solution and use this to create a bridge between your two beakers.
- Record the potential difference; if the reading is negative, swap the crocodile clips on the metals. Make a note of which metal is now connected to the negative terminal and so is the negative electrode.
- Repeat with different combinations of half-cells.

Safety

- Some materials are flammable.
- Wear goggles.
- Wash any spills on skin with plenty of water.

Experimental set-up



Worksheet G: Experimental analysis

Complete the following questions

The table below shows the literature values for the potential difference produced between magnesium, zinc, iron and copper half-cells.

Positive electrode	Negative electrode	Potential difference / V
zinc	magnesium	1.61
iron	magnesium	1.93
copper	magnesium	2.74
iron	zinc	0.32
copper	zinc	1.10
copper	iron	0.78

1. Are your results similar to the literature values for the half-cells used? Don't just think about values, think about patterns; is the biggest potential difference above for the same cell that produced the biggest potential difference you measured?
2. Why might your results be the same **OR** why might they be different?
3. Did any of your cells give an unexpected reading? If so, why might this have been?
4. Which electrode is always oxidised?
5. When completing the experiment did you notice any differences in the electrodes after the experiment was done which proves which electrode was oxidised?

Using the results from your experiment create a list of the half-cells from lowest electrode potential to the highest. Remember the negative electrode will have the more negative electrode potential.

Does this match the actual values? Why/Why not?

Worksheet H: Electrochemistry

Complete the following questions.

1. State the meaning of the term *standard electrode potential*.
2. Draw a diagram to show how you would set up an electrochemical cell.
3. What are standard conditions in relation to electrochemical cells?
4. A learner sets up an electrochemical cell with two vanadium-based half-cells. One is the V^{3+}(aq) | V^{2+}(aq) half-cell. The other involves a redox reaction between VO^{2+} and V^{3+}. The learner knows the electrode potential of the V^{3+}(aq) | V^{2+}(aq) half-cell is −0.26 V and that this is the negative electrode. The learner measures the potential difference of the cell to be 0.60 V. Calculate the electrode potential of the other half-cell.

The table below contains some standard electrode potential data. Use this data for the following questions.

Electrode half-equation	***E***$^{\ominus}$***/ V***
$F_2 + 2e^- \rightarrow 2F^-$	+2.87
$Au^+ + e^- \rightarrow Au$	+1.68
$2HOCl + 2H^+ + 2e^- \rightarrow Cl_2 + 2H_2O$	+1.64
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.23
$Ag^+ + e^- \rightarrow Ag$	+0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77
$2H^+ + 2e^- \rightarrow H_2$	0.00
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44

5. Identify the substance least likely to be oxidised in the table. Why is this the case?
6. Write the conventional representation of the half-cell involving Ag^+ and Ag.
7. A cell is made by connecting Fe^{2+} / Fe and Ag^+ / Ag electrodes with a salt bridge. What is the potential difference of this cell?
8. Suggest why potassium chloride would not be suitable for use in the salt bridge in the cell from question 7.
9. Nickel–cadmium cells are used to power electrical equipment such as drills and shavers.
   The electrode reactions are shown below.

   $$NiO(OH) + H_2O + e^- \rightarrow Ni(OH)_2 + OH^- \qquad E^{\ominus} = +0.52\ V$$

   $$Cd(OH)_2 + 2e^- \rightarrow Cd + 2OH^- \qquad E^{\ominus} = -0.88\ V$$

   (a) Calculate the potential difference of a nickel–cadmium cell.

   (b) Deduce an overall equation for the reaction that occurs in the cell when it is used.

10. A cell is made by connecting Al^{3+} / Al and Cu^{2+} / Cu electrodes with a salt bridge. Write the half-equations for the reactions that would occur.

Worksheet A: Answers

1. $Mg \rightarrow Mg^{2+} + 2e^-$
2. $Fe \rightarrow Fe^{2+} + 2e^-$
3. $2Cl^- \rightarrow Cl_2 + 2e^-$
4. $Cu^{2+} + 2e^- \rightarrow Cu$
5. $Ag^+ + e^- \rightarrow Ag$
6. $Fe \rightarrow Fe^{3+} + 3e^-$
7. $2O^{2-} \rightarrow O_2 + 4e^-$
8. $Al \rightarrow Al^{3+} + 3e^-$
9. $Fe^{2+} + 2e^- \rightarrow Fe$
10. $Zn \rightarrow Zn^{2+} + 2e^-$
11. $3Ag^+ + Fe \rightarrow Fe^{3+} + 3Ag$
12. $3Cu^{2+} + 2Al \rightarrow 3Cu + 2Al^3$

Worksheet B: Answers

1.

Zinc–copper electrochemical cell				
Electrode	**Charge of electrode**	**Half-equation**	**Oxidation or reduction?**	**Half-cell**
zinc	–	$Zn \rightarrow Zn^{2+} + 2e^-$	oxidation	$Zn^{2+}(aq) \mid Zn(s)$
copper	+	$Cu^{2+} + 2e^- \rightarrow Cu$	reduction	$Cu^{2+}(aq) \mid Cu(s)$

Magnesium–copper electrochemical cell				
Electrode	**Charge of electrode**	**Half-equation**	**Oxidation or reduction?**	**Half-cell**
magnesium	–	$Mg \rightarrow Mg^{2+} + 2e^-$	oxidation	$Mg^{2+}(aq) \mid Mg(s)$
copper	+	$Cu^{2+} + 2e^- \rightarrow Cu$	reduction	$Cu^{2+}(aq) \mid Cu(s)$

Silver–copper electrochemical cell				
Electrode	**Charge of electrode**	**Half-equation**	**Oxidation or reduction?**	**Half-cell**
silver	+	$Ag^+ + e^- \rightarrow Ag$	reduction	$Ag^+(aq) \mid Ag(s)$
copper	–	$Cu \rightarrow Cu^{2+} + 2e^-$	oxidation	$Cu^{2+}(aq) \mid Cu(s)$

2. The half-cell is always written so that the forward reaction shows reduction. The half-equations show the forward reaction at each half-cell so can be either oxidation or reduction.

3. More reactive metal is always oxidised.

4. Less reactive metal is always the positive electrode.

Worksheet C: Answers

Half-cell	$E^\ominus$ / V	Charge of electrode	Half-equation	Oxidation or reduction?
$Ag^+(aq)$ | $Ag(s)$	+0.80	+	$Ag^+ + e^- \rightarrow Ag$	reduction
$Fe^{2+}(aq)$ | $Fe(s)$	−0.44	−	$Fe \rightarrow Fe^{2+} + 2e^-$	oxidation

Half-cell	$E^\ominus$ / V	Charge of electrode	Half-equation	Oxidation or reduction?
$Mg^{2+}(aq)$ | $Mg(s)$	−2.37	−	$Mg \rightarrow Mg^{2+} + 2e^-$	oxidation
$Au^+(aq)$ | $Au(s)$	+1.68	+	$Au^+ + e^- \rightarrow Au$	reduction

Half-cell	$E^\ominus$ / V	Charge of electrode	Half-equation	Oxidation or reduction?
$Zn^{2+}(aq)$ | $Zn(s)$	+0.76	+	$Zn^{2+} + 2e^- \rightarrow Zn$	reduction
$Al^{3+}(aq)$ | $Al(s)$	−1.66	−	$Al \rightarrow Al^{3+} + 3e^-$	oxidation

Half-cell	$E^\ominus$ / V	Charge of electrode	Half-equation	Oxidation or reduction?
$Au^+(aq)$ | $Au(s)$	+1.68	+	$Au^+ + e^- \rightarrow Au$	reduction
$2H^+(aq)$ | $H_2(g)$	0.00	−	$H_2 \rightarrow 2H^+ + 2e^-$	oxidation

Half-cell	$E^\ominus$ / V	Charge of electrode	Half-equation	Oxidation or reduction?
$Fe^{2+}(aq)$ | $Fe(s)$	−0.44	−	$Fe \rightarrow Fe^{2+} + 2e^-$	oxidation
$Fe^{3+}(aq)$| $Fe^{2+}(aq)$	+0.77	+	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	reduction

Half-cell	$E^\ominus$ / V	Charge of electrode	Half-equation	Oxidation or reduction?
$Cl_2(g)$ | $Cl^-(aq)$	+1.36	+	$Cl_2 + 2e^- \rightarrow 2Cl^-$	reduction
$I_2(g)$ | $2I^-(aq)$	+0.54	−	$2I^- \rightarrow I_2 + 2e^-$	oxidation

Worksheet C: Answers, continued

Half-cell	**$E^\ominus$ / V**	**Charge of electrode**	**Half-equation**	**Oxidation or reduction?**
$Cl_2(g)$ | $Cl^-(aq)$	+1.36	−	$2Cl^- \rightarrow Cl_2 + 2e^-$	oxidation
$Au^+(aq)$ | $Au(s)$	+1.68	+	$Au^+ + e^- \rightarrow Au$	reduction

Half-cell	**$E^\ominus$ / V**	**Charge of electrode**	**Half-equation**	**Oxidation or reduction?**
$Cl_2(g)$ | $Cl^-(aq)$	+1.36	+	$Cl_2 + 2e^- \rightarrow 2Cl^-$	reduction
$2H^+(aq)$ | $H_2(g)$	0.00	−	$H_2 \rightarrow 2H^+ + 2e^-$	oxidation

Half-cell	**$E^\ominus$ / V**	**Charge of electrode**	**Half-equation**	**Oxidation or reduction?**
$Mg^{2+}(aq)$ | $Mg(s)$	−2.37	−	$Mg \rightarrow Mg^{2+} + 2e^-$	oxidation
$I_2(g)$ | $2I^-(aq)$	+0.54	+	$I_2 + 2e^- \rightarrow 2I^-$	reduction

The voltmeter reading is the more positive number take away the less positive number

$$0.34 - (-0.76)$$

This is how we work out the cell potential.

Worksheet D: Answers

1.
   a. $+0.34 - (-2.37) = 2.71$ V
   b. $+0.80 - (-2.37) = 3.17$ V
   c. $-0.44 - (-1.66) = 1.22$ V
   d. $+0.52 - (-0.14) = 0.66$ V
   e. $-0.14 - (-0.76) = 0.62$ V
   f. $+1.09 - (+0.34) = 0.75$ V
   g. $+1.98 - (-0.44) = 2.42$ V
   h. $+1.09 - (-1.66) = 2.75$ V
   i. $+0.52 - (+0.34) = 0.18$ V
   j. $-0.14 - (-2.37) = 2.23$ V

2.
   a. Yes.
   b. No, as copper is being oxidised, as it has the more positive number it should be being reduced.
   c. No, iron should be oxidised as less positive number than bromine.
   d. Yes.

3.
   a. There will be a reaction as Pb can be oxidised and H^+ reduced.

   $$Pb + 2H^+ \rightarrow Pb^{2+} + H_2$$

   b. No reaction

Worksheet E: Answers

There are no wrong answers for this worksheet, the idea is to come up with things the learners see as missing or incorrect in the method.

Possible problems students may come up with are listed below.

<u>Equipment needed</u>
- Missing crocodile clips and wires.
- Doesn't state the concentration of the salt solutions (would need to be 1 mol dm^{-3} for standard conditions.

<u>Method</u>
- Not specific, doesn't say the salt solution and metal rod need to be the same metal.
- Doesn't say the salt bridge needs to be soaked in salt solution.
- Doesn't mention what to do if the potential difference is negative.

<u>Safety</u>
- Doesn't say what to do to address any problems. Materials are flammable so keep them away from fire. Some salt solutions are irritants so wear goggles and rinse any spills on skin with plenty of water.

<u>Errors</u>
- Doesn't say how to fix the error. Read the potential difference immediately after salt bridge is placed, as an example.

<u>Diagram of experimental setup</u>
- Could have more labels

Worksheet G: Answers

The answers for this worksheet will be based on the learners own results.

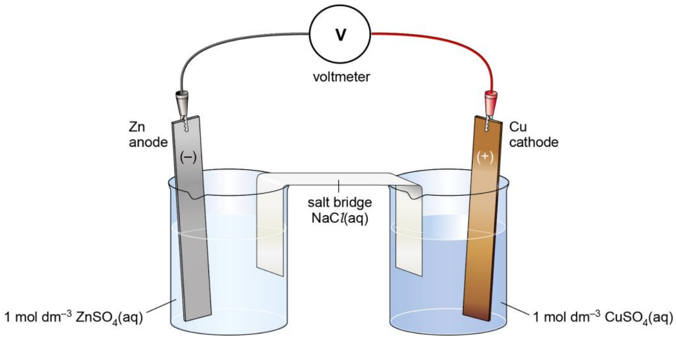
1. This answer will depend on the learner's practical results. Answers could relate to which metals are always the negative electrode (e.g. magnesium as it is the most reactive). The biggest potential difference should be caused by the most reactive and least reactive metal together (Mg and Cu in this case) and the smallest by those metals with the least difference in reactivity (Fe and Zn in this case).
2. The values in the literature are based on standard conditions. Learners could talk about the conditions during the practical being slightly different from standard, e.g. temperature of the room, slightly different concentrations of the solutions (particularly if they used the same solution at the end of the practical as at the start and they had been left in an electrochemical cell for a while). They could also talk about the idea that the metal rods became tarnished.
3. This answer will depend on learner's results. Answer could be yes or no. If yes, differences could have been caused by non-standard conditions – e.g. temperature of the room, slightly different concentrations of the solutions (particularly if they used the same solution at the end of the practical as at the start and they had been left in an electrochemical cell for a while). They could also talk about the idea that the metal rods became tarnished.
4. The negative electrode.
5. The negative electrode should have degraded slightly and the metal becomes metal ions. The positive electrode should have a fresh layer of the metal on it.

Again, this will depend on their results, but answer should be Mg, Zn, Fe, Cu

Any differences could have been caused by non-standard conditions.

Worksheet H: Answers

1. The potential difference of the half-cell connected to a standard hydrogen half-cell.
2.



3. Salt concentration of 1 mol dm^{-3}, 25 °C, 1 atm pressure.
4. 0.60 + (–0.26) = +0.34 V
5. Fluorine as it has the most positive value, which means it is more reactive at gaining electrons and so will be reduced.
6. $Ag^{+}(aq) \mid Ag(s)$
7. +1.24 V
8. The chloride ions would react with the silver ions to make the precipitate silver chloride.
9. (a) +1.40 V

   (b) $2NiO(OH) + 2H_2O + Cd \rightarrow 2Ni(OH)_2 + Cd(OH)_2$

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

    $Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$

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