ROYAL SOCIETY
OF **CHEMISTRY**

Electrochemical cells

Practical video

Supporting resources

Contents

Teacher notes

This resource supports the practical video Electrochemical cells, available here: rsc.li/3l0g6sR

The value of experiencing live practical work cannot be overstated. Numerous studies provide evidence of its value in terms of learner engagement, understanding, results and the likelihood of continuing to study chemistry or work in a related field. This video can be used to complement live practical work, as well as helping learners to understand the methods, equipment and skills when they cannot access the lab.

How to use this video

The video and additional resources are designed to be used flexibly, but some suggestions follow.

Flipped learning

Learners view the video ahead of the live practical lesson to help it run more smoothly and keep objectives in focus. This may also help build confidence for some learners and improve their outcomes in the lesson. Use questions from the pause-and-think set provided as part of the preparation task.

Consolidation and revision

Learners view the video after the practical – this may be directly after the lesson or learners can return to it as part of revision for examinations.

Revisiting a practical with a different focus

A practical experiment can support many learning outcomes. Focussing on just one or two of those in a lesson will help ensure that the aims are achieved. The video could be used to revisit the experiment with a different focus.

Home learning

Whether it is remote teaching, homework, or individual learner absence, the video provides an opportunity to engage with a practical experiment and the associated skills when learners are not in the lab.

Other tips

- **Provide your own commentary**
  Mute the voice over and provide your own commentary. This will allow you to better engage with learners and adapt to the needs and objectives of your lesson.

- **Use questions**
  A set of pause-and-think questions are provided in two formats, one for teacher-led questions and discussion and a student worksheet which can be used independently by learners. Select from these or create your own questions to help engage learners and target specific aims.

Notes on running the practical experiment

Technician notes including the equipment list and safety notes are available as a separate document here: rsc.li/3l0g6sR. If you are planning to carry out the practical, you will need to carry out your own risk assessment.

Instructions for learners

A student sheet with instructions for carrying out the experiment is available here: rsc.li/3l0g6sR

Results tables

Printable and editable results tables for the two microscale experiments shown in the video are provided.

Further practical activities

The demonstration The oxidation states of vanadium explores different colours of different oxidation states of the transition metals and continues to develop the knowledge that electrode potentials can be used to help predict the course of redox reactions.

Our exhibition chemistry video Nailing corrosion demonstrations looks at rusting from an electrochemical perspective.

Prior knowledge

The video and resources assume the following prior knowledge:

- Metals can be put into a reactivity series.

- The more reactive metal is a stronger reducing agent.

- Zinc is more reactive than copper.

- Oxidation and reduction defined in terms of electron transfer - OIL RIG.

- Electrons are negatively charged and move from a negative terminal to a positive terminal.

- ΔG is negative for feasible reactions.

Learners could be directed to our 14–16 video with supporting resources Electrolysis of aqueous solutions, for revision prior to this practical.

Common misconceptions

Be aware of misconceptions learners may have around this practical, for example:

- Only metals can be placed in the reactivity series.

- Large quantities of solutions are needed.

- A sizeable current is flowing when we measure the voltage.

- How the salt bridge works.

- Surface area of the metal will change the electrode potential.

This article, Redox: reducing difficulties, looks at common misconceptions more closely and suggests strategies to address them.

Intended outcomes

It is important that the purpose of each practical is clear from the outset, defining the intended learning outcomes helps to consolidate this. Outcomes can be categorised as hands on, what learners are going to do with objects, and minds on, what learners are going to do with ideas to show their understanding. We have offered some differentiated suggestions for this practical. You may wish to focus on just one or two, or make amendments based on your learners' own needs. (Read more at rsc.li/2JMvKa5.)

Consider how you can share outcomes and evaluations with learners, empowering them to direct their own learning

	Hands on	Minds on
Effective at a lower level	**Learners correctly:** • set up a functioning electrochemical cell • swap the connectors round to change the potential difference from negative to positive • accurately measure and record the potential difference and which metal is connected to each terminal	**Learners can:** • understand that the voltage is related to the relative reactivity of the metals • understand that the more reactive metal will have the more negative potential • recall and apply OIL RIG to negative and positive ends of the cell • recall that changing the concentration of solution changes the potential • understand that other conditions such as temperature may affect the cell potential
Effective at a higher level	**Learners correctly:** • predict which half-cell will be positive and connect the voltmeter accordingly when dealing with common metals in the reactivity series • build accurate cell diagrams for the cells tested	**Learners can:** • understand that the cells work through forcing the electrons round the external circuit • use a measured cell potential to predict a feasible reaction • apply Le Chatelier's principle to explain the effect of changing concentration of solution • understand the greater the voltage the greater the free energy change for the feasible reaction

How to use the additional resources

Using the pause-and-think questions

Pause-and-think questions are supplied in two formats: a teacher version for 'live' questioning and a student version which can be used during independent study. The time stamps allow you to pause the video when presenting to a class, or learners to use for active revision.

Teacher version

The questions are presented in a table and you can choose to use as many as appropriate for your class and the learning objectives.

Some questions have two timestamps to allow you to adapt the questions for different classes or scenarios. Pause the videos at the earlier timestamp to ask a question before the answer is given, useful for revision or to challenge learners. Pause at the later timestamp to ask a question reflectively and assess whether learners have understood what they have just heard or seen. This would be useful when introducing a topic, in a flipped learning scenario or when additional support and encouragement is needed.

Think about how you will ask for responses. Variation may help to increase engagement – learners could write and hold up short answers; more complex questions could be discussed in groups.

Not all answers to questions are included in the video. Some of the questions will draw on prior learning or extend learners' thinking beyond the video content.

Student version

The same questions are offered as a printable worksheet for learners. Use in situations where there is not a teacher present to guide discussion during the video, for example homework, revision or remote learning.

Using the follow-up worksheet

A follow-up worksheet has been included as part of these additional resources. This worksheet could be used to follow up the practical activity, for example as homework or a revision exercise.

Additional resources

Pause-and-think questions

Teacher version

Timestamp(s)	Question	Answer/discussion points
00:34/00:38	What does the mnemonic OIL RIG stand for?	**O**xidation **i**s **L**oss of electrons. **R**eduction **i**s **G**ain of electrons.
01:35/01:53	How do electrolytes conduct electricity?	Electrolytes are composed of free moving charged particles, ions. Negative ions are attracted to the positive electrode where they lose electrons; positive ions are attracted to the negative electrode where they gain electrons. It is this movement of ions in the solution that conducts electricity.
02:43	Can you place copper, zinc and magnesium in order of reactivity?	Magnesium > zinc > copper (from most reactive to least reactive).
04:10	Zinc is higher in the reactivity series than copper. Why does it have a more negative potential than copper?	The zinc metal lattice gives up electrons more easily, its atoms are more readily oxidised than copper atoms so electrons flow from the zinc electrode to the copper.
04:36/04:41	What voltage would we expect if we swapped over the leads connected to the zinc and the copper?	It should be equal in magnitude but opposite in sign.
04:41	Which of these two reactions does the cell potential suggest will happen: a. zinc reducing copper ions, b. copper reducing zinc ions?	Zinc reducing copper ions as electrons flow from the zinc electrode to the copper.
05:18	Record the potential difference and which metal was connected to the negative terminal and which to the positive terminal to give a positive value.	0.62 V The negative terminal was connected to the magnesium and the positive terminal was connected to the zinc.
05:26	What is the relationship between the distance apart in the reactivities series of the metals and the potential difference between their half-cells?	The greater the difference in relative reactivity of the metals the greater the potential difference.

Timestamp(s)	Question	Answer/discussion points
05:56/06:04	We measured the potential difference between the copper and zinc to be 0.96 V. The potential difference between copper and iron is 0.50 V. Can we state which metal, iron or zinc, is higher in the reactivity series and predict what the potential difference between iron and zinc will be?	Zinc must be higher in the reactivity series because the copper/zinc cell has a greater potential than the copper/iron cell. Both the iron and zinc were the negative terminals of the cell showing that they were more reactive than copper. It therefore seems reasonable to suggest the potential difference will be the difference 0.96 – 0.50 = 0.46 V and indeed it is. (It was 0.44 V in the video, but the voltmeter had an uncertainty of +/- 0.02 V.) However, this is only the case where the number of electrons in the half equations is the same. The electrode potentials give us an indication of ΔG for the reactions and ΔG is proportional to the cell potential difference multiplied by the number of electrons in the half equations. We can illustrate this point with some reduction potentials of gold:

Electrode reaction	$E^\oplus$/V
$Au^+ + e^- \rightleftharpoons Au$	+1.68
$Au^{3+} + 2e^- \rightleftharpoons Au^+$	+1.41
$Au^{3+} + 3e^- \rightleftharpoons Au$	+1.50

Learners might have expected the Au/Au^+ and Au^+/Au^{3+} potentials to add together to equal the Au^+/Au^{3+} potential.
Note that 3 x 1.5 = 2 x 1.41 + 1.68.
They do add up if we take into account the number of electrons in each half equation. |
07:36	Draw a cell diagram for the Zn/Mg cell.	$Mg(s) \mid Mg^{2+}(aq) \parallel Zn^{2+}(aq) \mid Zn(s)$
08:33/08:38	Two of the concentrations are the same, 1.0 mol dm^{-3}. What will the potential difference be between these two half-cells?	It should be 0 V as the two half-cells are identical and therefore the equilibriums in the half-cells should be equivalent.
08:48/08:57	What do you think will happen when the electrodes are applied to two different concentrations?	A potential difference will be recorded.

Timestamp(s)	Question	Answer/discussion points
09:47	Now we have seen that concentration of solutions is one condition that can affect the potential difference of a half-cell, can you think of any other variables that could affect the potential difference of a half-cell? How might you investigate them?	Temperature affects the cell potential and also pressure if gases are involved, but not surface area of the metal. Temperature could be investigated in this microscale set-up by changing the temperature in one half of the cell, eg standing the petri dish with an ice cube under one half and a room temperature support under the other. Another method would be to set up the apparatus except for the solution in one half-cell, which is heated before adding. The traditional beaker method is likely to be easier to investigate temperature as a larger volume of solution is more easily controlled. It can also be more obvious to learners if you demo this using two copper half-cells identical in every other respect, so they can see the voltage change as one half-cell is cooled or heated.
09:48	Can you use Le Chatelier's principle to explain why lowering the concentration of the copper sulphate solution made the half-cell more negative?	If we consider the half equation as an equilibrium $Cu^{2+} + 2e^- \rightleftharpoons Cu$, then if we lower the concentration of $Cu^{2+}(aq)$ Le Chatelier's principle would predict a shift in the position of equilibrium to the left and more oxidation of copper atoms to ions.
10:06	Referring to the copper/zinc reaction we saw at the start of the video. If more zinc ions are being formed in the zinc half-cell, which ions in the salt bridge will flow into the zinc half-cell? If the copper ions are being removed from the copper solution, which ions from the salt bridge flow to that side? Use these ideas to explain how salt bridges work.	The solutions in each half-cell should remain neutrally charged. If we are forming positive zinc ions in the zinc half-cell that would suggest a flow of negative ions in the zinc half-cell and/or a flow of positive ions (Zn^{2+}) out of the zinc half-cell. The removal of positive ions from the copper half-cell would suggest a flow of positive ions from the salt bridge into the copper half-cell and/or a flow of negative ions out of the half-cell. The salt bridge then allows the flow of ions in and out of each half-cell so the solutions can remain neutral. It is worth noting that the use of a high resistance voltmeter means the actual flow of current is very small – just enough to establish the potential so the flow of ions does not have to be either large or sustained which is why a piece of filter paper soaked in appropriate salt will do.

Pause-and-think questions

Student version

Pause the video at the time stated to test or revise your knowledge of these practical experiments.

Time	Question
00:34	What does the mnemonic OIL RIG stand for?
02:43	Can you place copper, zinc and magnesium in order of reactivity?
03:11	How do electrolytes conduct electricity?
04:10	Zinc is higher in the reactivity series than copper. Why does it have a more negative potential than copper?
04:38	What voltage would we expect if we swapped over the leads connected to the zinc and the copper?
04:41	Which of these two reactions does the cell potential suggest will happen:

a. zinc reducing copper ions,

b. copper reducing zinc ions? |
05:18	Record the potential difference and which metal was connected to the negative terminal and which to the positive terminal to give a positive value.
05:26	What is the relationship between the distance apart in the reactivities series of the metals and the potential difference between their half-cells?
05:56	We measured the potential difference between the copper and zinc to be 0.96 V. The potential difference between copper and iron is 0.50 V. Can we state which metal, iron or zinc, is higher in the reactivity series and predict what the potential difference between iron and zinc will be?
07:36	Draw a cell diagram for the Zn/Mg cell.
08:33	Two of the concentrations are the same, 1.0 mol dm^{-3}. What will the potential difference be between these two half-cells?
08:48	What do you think will happen when the electrodes are applied to two different concentrations?
09:47	Now we have seen that concentration of solutions is one condition that can affect the potential difference of a half-cell, can you think of any other variables that could affect the potential difference of a half-cell? How might you investigate them?
09:48	Can you use Le Chatelier's principle to explain why lowering the concentration of the copper sulphate solution made the half-cell more negative?
10:06	Referring to the copper/zinc reaction we saw at the start of the video. If more zinc ions are being formed in the zinc half-cell, which ions in the salt bridge will flow into the zinc half-cell? If the copper ions are being removed from the copper solution, which ions from the salt bridge flow to that side? Use these ideas to explain how salt bridges work.

Follow-up worksheet

In the video Electrochemical cells, available here: rsc.li/3l0g6sR, a series of microscale cells were set-up and tested. The results are given in the table below:

Half-cell combination	Which metal formed the negative terminal?	Potential difference of the cell (V)
Zn/Cu	zinc	±0.96
Zn/Mg	magnesium	±0.62
Fe/Mg	magnesium	±1.07
Cu/Fe	iron	±0.50
Fe/Zn	zinc	±0.44
Cu/Mg	magnesium	±1.55

1. What is the order of reactivity of the four metals tested?

2. Does oxidation or reduction occur at the half-cell with the negative electrode?

3. Predict the feasible reaction occurring in the Zn/Mg cell.

4. Draw a cell diagram for the Zn/Mg cell.

5. Predict the feasible reaction occurring in the Zn/Fe cell.

6. Use your results to decide whether each of the following reactions is feasible or non-feasible:
   a. $Mg^{2+}(aq) + Fe(s) \rightarrow Mg(s) + Fe^{2+}(aq)$
   b. $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$
   c. $Zn^{2+}(aq) + Fe(s) \rightarrow Zn(s) + Fe^{2+}(aq)$
   d. $Mg(s) + Fe^{2+}(s) \rightarrow Mg^{2+}(aq) + Fe(s)$
   e. $Mg(s) + Zn^{2+}(s) \rightarrow Mg^{2+}(aq) + Zn(s)$

7. We tested another metal M and set up a cell with the iron half-cell and a M/M^{2+} half-cell then measure a potential difference of +-.19 V with iron as the negative half-cell.
   a. Put metal M into your order of reactivity from question 1.
   b. Write an equation for the feasible reaction between M(s) or $M^{2+}(aq)$ and Fe(s) or $Fe^{2+}(aq)$.

Follow-up worksheet: answers

1. What is the order of reactivity of the four metals tested?

   Most reactive to least reactive: magnesium > zinc > iron > copper

2. Does oxidation or reduction occur at the half-cell with the negative electrode?

   Oxidation, since electrons flow out from the negative half-cell something is losing electrons therefore is being oxidised.

3. Predict the feasible reaction occurring in the Zn/Mg cell.

   Because the zinc half-cell is the positive half-cell reduction is occurring ($Zn^{2+} + 2e^- \rightarrow Zn$) and the magnesium half-cell is negative so oxidation is occurring there.

4. Draw a cell diagram for the Zn/Mg cell.

   Mg(s) I Mg^{2+}(aq) II Zn^{2+}(aq) I Zn(s)

5. Predict the feasible reaction occurring in the Zn/Fe cell.

   Zn(s) + Fe^{2+}(aq) → Fe(s) + Zn^{2+}(aq)

6. Use your results to decide whether each of the following reactions is feasible or non-feasible:

   a. Mg^{2+}(aq) + Fe(s) → Mg(s) + Fe^{2+}(aq)

   Non-feasible (as the magnesium was the negative half-cell of these two)

   b. Cu^{2+}(aq) + Fe(s) → Cu(s) + Fe^{2+}(aq)

   Feasible (as the copper ions are reduced and the copper half-cell is indeed the more positive of these two)

   c. Zn^{2+}(aq) + Fe(s) → Zn(s) + Fe^{2+}(aq)

   Non-feasible (for the zinc to be reduced by iron the zinc would need to be the more positive of the two half-cells)

   d. Mg(s) + Fe^{2+}(s) → Mg^{2+}(aq) + Fe(s)

   Feasible (as the iron(II) ions are reduced and the iron half-cell is indeed the more positive of these two)

   e. Mg(s) + Zn^{2+}(s) → Mg^{2+}(aq) + Zn(s)

   Feasible (as Mg is the more negative of the two half-cells so can be oxidised)

7. We tested another metal M and set up a cell with the iron half-cell and a M/M^{2+} half-cell then measure a potential difference of +-.19 V with iron as the negative half-cell.

   a. Put metal M into your order of reactivity from question 1.

   Most reactive to least reactive: magnesium > zinc > iron > 'M' > copper

   b. Write an equation for the feasible reaction between M(s) or M^{2+}(aq) and Fe(s) or Fe^{2+}(aq).

   M^{2+}(aq) + Fe(s) → M(s) + Fe^{2+}(aq) as the more negative half-cell (in this case the iron) will be where the oxidation is taking place.

Printable results tables

Half-cell combination	Which metal formed the negative terminal?	Potential difference of the cell (V)
Zn/Cu		
Zn/Mg		
Fe/Mg		
Cu/Fe		
Fe/Zn		
Cu/Mg		

Half-cell combination	Which metal formed the negative terminal?	Potential difference of the cell (V)
Zn/Cu		
Zn/Mg		
Fe/Mg		
Cu/Fe		
Fe/Zn		
Cu/Mg		

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Zn/Cu		
Zn/Mg		
Fe/Mg		
Cu/Fe		
Fe/Zn		
Cu/Mg		

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Zn/Cu		
Zn/Mg		
Fe/Mg		
Cu/Fe		
Fe/Zn		
Cu/Mg		

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Zn/Cu		
Zn/Mg		
Fe/Mg		
Cu/Fe		
Fe/Zn		
Cu/Mg		

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Zn/Cu		
Zn/Mg		
Fe/Mg		
Cu/Fe		
Fe/Zn		
Cu/Mg		

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Zn/Mg		
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Fe/Zn		
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Cu/Mg		