Education in Chemistry 16–18 years
Available from rsc.li/3XrKEXh

Electrochemical cells misconception buster

This resource accompanies the article **Teaching electrochemical cells at post-16** in *Education in Chemistry* which can be viewed at: rsc.li/46nHSXn. The article provides classroom tips, activities and ideas to combat electrochemistry misconceptions.

Learning objectives

1. Describe how to set up an electrochemical cell, including:
    - The function and use of a salt bridge.
    - The relative positions of half cells according to their $E^\ominus$ values.
    - The use of a platinum electrode where necessary.
    - The components and use of the standard hydrogen electrode.
2. Write and apply the conventional representation of an electrochemical cell.
3. Use $E^\ominus$ values to predict the direction of simple redox reactions.
4. Calculate the EMF (E_{cell}) and use this value to predict the feasibility of the cell.

Identify and target misconceptions in the three main pinch points of electrochemistry: setting up cells, redox equations and calculations.

How to use the resource

Activity 1 consists of 20 multiple choice questions. Instruct learners to read each question carefully, using the diagrams and data provided to tick an answer. There is one correct answer per question.

After learners mark the multiple choice questions either by peer or self-assessing, they will complete the relevant follow-up tasks to target areas for improvement and increase their understanding of certain topics.

If appropriate, guide learners to follow-up tasks based on misconceptions you have already identified throughout the course or for further practice if they get 20/20 on the quiz. Activity 2 is fully differentiated for learners to target their own misconceptions identified from the multiple choice quiz.

Extension opportunities

Non-rechargeable and rechargeable cells are not explicitly covered in this task. Encourage learners to read the *EiC* **Batteries and electrochemical cells** article (available from rsc.li/3Xoy23g) as an introduction. As a follow-up activity, use the lead–acid accumulator cell practical (rsc.li/3NtKAC0) to either introduce or revise these concepts.

Answers

Activity 1: multiple choice questions

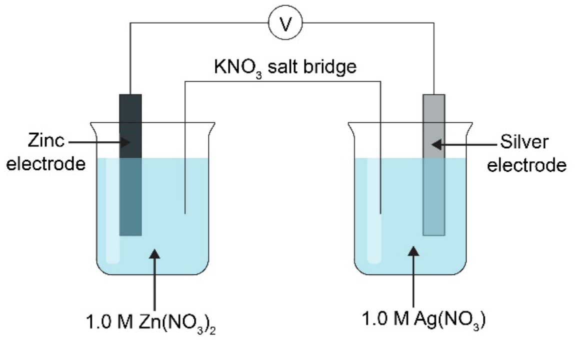
Qu	Answer	Explanation	Activity 2 follow-up task
1	A – A half-cell	An electrochemical cell is made up of two half-cells joined together by a connecting wire and a salt-bridge. A battery is made of **multiple** cells joined together in series.	Complete **task a**
2	C – $0.5\ \text{mol dm}^{-3}$	The concentration of **hydrogen ions** in the standard hydrogen electrode must be $1.00\ \text{mol dm}^{-3}$. Since sulfuric acid is a **diprotic acid** with the formula H_2SO_4, every mole of acid releases two moles of hydrogen ions, therefore the concentration of acid must be $0.5\ \text{mol dm}^{-3}$ to release $1.00\ \text{mol dm}^{-3}$ of hydrogen ions.	Complete **task b**
3	D - $Fe^{2+}(aq), Fe^{3+}(aq)\|Pt(s)$	According to IUPAC convention, reduced forms go on the outermost edge of cell representations. You separate species in the same phase by a comma (,), while separating species in a different phase by a single line (|). Since Fe^{2+} and Fe^{3+} are both aqueous, you use an inert platinum electrode.	Complete **task c**
4	A - Reduced form at edges, more negative cell on the left-hand side.	According to IUPAC convention, reduced forms go on the outermost edge of the cell representation. The more negative half-cell goes on the left.	Complete **task c**
5	C - 298 K, 100 kPa, $1\ \text{mol dm}^{-3}\ H^+$ solution, platinum electrode.	The concentration of **hydrogen ions** in the standard hydrogen electrode (SHE) must be $1.00\ \text{mol dm}^{-3}$, not simply the 'acid' concentration (as some acids, such as sulfuric acid, release more than one hydrogen ion per acid molecule). You use	Complete **task b**

		standard conditions for the SHE. A graphite electrode is unsuitable, as it may react with the solutions; you must use an inert platinum electrode.	
6	B - It allows inert ions to transfer between half-cells to maintain electrical neutrality.	The salt bridge completes the electrical circuit as described. Note that **only** the inert ions used in the salt bridge can transfer between solutions; **not** the ions from the electrochemical cell. Neither current nor electrons can 'flow' through a salt bridge.	Complete **task a**
7	C – KNO_3	Use salts which do not easily form precipitates for salt bridges. Sodium ions and chloride ions are inappropriate, as they are too large to allow effective diffusion across the salt bridge. Chloride forms many insoluble salts with metals, so cannot be used for salt bridges.	Complete **task a**
8	A - | |	You must represent a salt bridge with a double vertical line, | |.	Complete **task a**
9	C – $Pt(s)$	To construct the SHE, pump hydrogen gas through a porous tube containing an inert platinum electrode dipped in a $1\ mol\ dm^{-3}$ H^+ solution at $298\ K$ and $100\ kPa$.	Complete **task b**
10	D – Copper electrode	To construct a half-cell, dip a metal into a solution of its ions. Since the solution is copper sulfate, you must use a copper electrode. Only use a platinum electrode if there are no solid electrodes in the half-cell.	Complete **task a**
11	D – The cell with the more negative $E^{\ominus}$ value.	Half-cells can have positive or negative $E^{\ominus}$ values, depending on the tendency of the species to be oxidised or reduced. The **more** negative cell is always placed on the left-hand side; it is the anode.	Complete **task a**
12	A – Electrons flow from the zinc half-cell to the iron half-cell; Zn is oxidised and Fe^{3+} is reduced.	The zinc loses electrons to form Zn^{2+}. The electrons flow through the conducting wire to the iron half-cell. Fe^{3+} ions gain an electron to form Fe^{2+}.	Complete **task d**

13	A – $\mathrm{Zn(s)\|Zn^{2+}(aq)\|\|Fe^{3+}, Fe^{2+}(aq)\|Pt(s)}$	According to IUPAC convention, reduced forms go on the outermost edge of the cell representation, so Fe^{2+} must go on the far right, and Zn on the far left. The more negative half-cell goes on the left. Since there are no solid electrodes in the Fe^{2+}/Fe^{3+} half-cell, you require a platinum electrode.	Complete **task a**
14	C – Cl_2	Oxidising agents cause other species to become oxidised. They themselves are reduced.	Complete **task f**
15	B – Br^-	The bromine/bromide half-cell has a more-negative (less positive) $E^{\ominus}$ value, so oxidation occurs, and bromide ions lose electrons to form bromine: $2Br^- \rightleftharpoons Br_2 + 2e^-$	Complete **task d**
16	A – $Cl_2 + 2Br^- \rightleftharpoons 2Cl^- + Br_2$	The chlorine/chloride half-cell has a more-positive $E^{\ominus}$ value, so reduction occurs. $Cl_2 + 2e^- \rightleftharpoons 2Cl^-$ This means that the bromine/bromide reaction occurs in the opposite direction. $2Br^- \rightleftharpoons Br_2 + 2e^-$ Combining the equations gives A.	Complete **task e**
17	A – Cl_2/Cl^- as electrons flow towards it	Electrons flow **towards** the positive electrode. Since the $E^{\ominus}$ value for the chlorine/chloride half-cell is more-positive, we know that reduction occurs there – electrons are gained. So, electrons must flow **towards** this half-cell.	Complete **task f**
18	B – $+0.27\ V$	$E_{cell} = E^{\ominus}_{red} - E^{\ominus}_{ox}$ We know that the chlorine/chloride half-cell is where reduction occurs. $+1.36 - 1.09 = +0.27\ V$	Complete **task g**
19	D - Mg	Reducing agents cause other species to become reduced. They themselves are oxidised. Species with a higher tendency to become oxidised have a more-negative $E^{\ominus}$ value. Ensure you are selecting the species which loses electrons – the Mg loses two	Complete **task f**

		electrons to become Mg^{2+}, so it is Mg which is the reducing agent.	
20	C – ii and iv only	$E_{cell} = E^{\ominus}_{red} - E^{\ominus}_{ox}$ Reactions are only feasible if they produce a positive EMF (E_{cell}). The EMFs for the cells as shown are: i. $-0.25 - 0.34 = -0.59$ V ii. $+0.77 - (-2.33) = 3.10$ V iii. $-0.25 - (0.77) = -1.02$ V iv. $+0.34 - (-0.25) = +0.59$ V	Complete **task h**

Activity 2: follow-up tasks

Follow-up task	Answer
(a)	Credit labelled diagrams or written descriptions. Zinc half-cell – 1.00 mol dm^{-3} of Zn^{2+} ions (allow any suitable solution such as $ZnSO_4$, $Zn(NO_3)_2$, $ZnCl_2$ etc). $Zn(s)$ electrode submerged into the solution, connected to a conducting wire. Silver half-cell – 1.00 mol dm^{-3} of Ag^{+} ions (allow any suitable solution such as $AgNO_3$, but NOT silver halide solutions as they form precipitates). $Ag(s)$ electrode submerged into the solution, connected to a conducting wire. Salt bridge connects the two cells to complete the circuit. Any suitable ions – eg KNO_3. Zinc half-cell must be placed on the left-hand side. Voltmeter used in the circuit to measure the EMF. 
(b)	Credit labelled diagrams or written descriptions. Standard hydrogen electrode (SHE): 1.00 mol dm^{-3} H^{+} solution (check the solution used – if $H_2SO_4(aq)$, then the concentration of acid must be 0.5 mol dm^{-3}). Platinum electrode submerged in the acid solution, connected to a conducting wire. $H_2(g)$ pumped through a porous tube into the solution.

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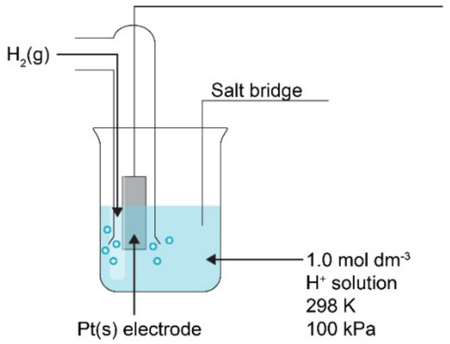
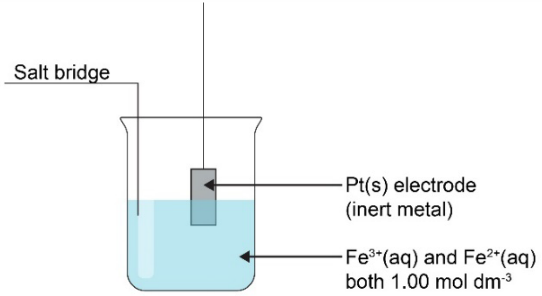
SHE placed on the left-hand side of the electrochemical cell, as shown in the diagram:



Fe^{2+}/Fe^{3+} half-cell:
1.0 mol dm^{-3} both Fe^{2+} and Fe^{3+} solution.
Pt(s) electrode connected to a conducting wire.
As shown in the diagram:



Salt bridge connects the two cells to complete the circuit. Any suitable ions – eg KNO_3.
SHE must be placed on the left-hand side.
Voltmeter used in the circuit to measure the EMF.
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- The more negative cell, with the exception of the SHE, which is always placed on the left.
- Any cells with non-solid electrodes. Common examples include the Fe^{2+}/Fe^{3+} half-cell, and non-metal half-cells such as Cl_2/Cl.
- Platinum is inert; graphite could react with the electrolyte solutions.
- SHE is always placed on the left-hand side of a cell.
- The right-hand side.
- From the anode (LHS) to the cathode (RHS), ie from the more-negative cell to the less-negative cell.
- Reduced forms go on the edges of the cell. Substances in a different phase are separated by a single vertical line, | . Substances in the same phase are separated by a comma. The salt bridge is represented by a double vertical line, | | .

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$Cl_2 + 2e^- \rightleftharpoons 2Cl^- \quad E^\ominus = +\ 1.36\ V$
$Br_2 + 2e^- \rightleftharpoons 2Br^- \quad E^\ominus = +1.09\ V$
$2Br^- + Cl_2 \rightleftharpoons 2Cl^- + Br_2$
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	The Br_2/Br^- half-cell is more-negative: oxidation occurs. Electrons flow from the Br_2/Br^- half-cell (anode) to the Cl_2/Cl^- half-cell (cathode). The reaction is feasible in this direction since the EMF would be positive: $\text{EMF} = +1.36 - 1.09 = +0.27\ \text{V}$
(e)	Reaction ii: Mg^{2+}/Mg half-cell is more-negative – Mg must be oxidised to Mg^{2+}: $Mg \rightleftharpoons Mg^{2+} + 2e^-$ Fe^{2+}/Fe^{3+} half-cell is more-positive – Fe^{3+} must be reduced to Fe^{2+}: $Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$ Combining the two half-equations gives the ionic equation: $Mg + 2Fe^{3+} \rightleftharpoons Mg^{2+} + 2Fe^{2+}$ Reaction iv: Ni^{2+}/Ni half-cell is more-negative – Ni must be oxidised to Ni^{2+}: $Ni \rightleftharpoons Ni^{2+} + 2e-$ Cu^{2+}/Cu half-cell is more-positive – Cu^{2+} must be reduced to Cu: $Cu^{2+} + 2e^- \rightleftharpoons Cu$ Combining the two half-equations gives the ionic equation: $Cu^{2+} + Ni \rightleftharpoons Ni^{2+} + Cu$ (This is also the feasible direction of reaction i). Feasible reaction iii: Fe^{2+}/Fe^{3+} half-cell is more-positive – Fe^{3+} must be reduced to Fe^{2+}: $Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$ Ni^{2+}/Ni half-cell is more-negative – Ni must be oxidised to Ni^{2+}: $Ni \rightleftharpoons Ni^{2+} + 2e-$ Combining the two half-equations gives the ionic equation: $Ni + 2Fe^{3+} \rightleftharpoons Ni^{2+} + 2Fe^{2+}$
(f)	• Oxidation – loss of electrons. • Reducing agent – a species which causes other species to be reduced by gaining electrons; it is oxidised (loses electrons) in the process. • Cathode – the electrode where reduction occurs. $Mg + 2Fe^{3+} \rightleftharpoons Mg^{2+} + 2Fe^{2+}$ Mg is the reducing agent (since it is oxidised to Mg^{2+}). Fe^{3+} is the oxidising agent (since it is reduced to Fe^{2+}). Mg^{2+}/Mg electrode is the anode (oxidation occurs). Fe^{2+}/Fe^{3+} electrode is the cathode (reduction occurs).

	$Cu^{2+} + Ni \rightleftharpoons Ni^{2+} + Cu$ Ni is the reducing agent (since it is oxidised to Ni^{2+}). Cu^{2+} is the oxidising agent (since it is reduced to Cu). Ni^{2+}/Ni electrode is the anode (oxidation occurs). Cu^{2+}/Cu electrode is the cathode (reduction occurs).
(g)	$EMF = +0.80 - (-2.38) = +3.18\ V$
(h)	Feasible direction is: $Cr^{2+} \rightleftharpoons Cr^{3+} + e^-$ $Pb^{2+} + 2e^- \rightleftharpoons Pb$ $2Cr^{2+} + Pb^{2+} \rightleftharpoons 2Cr^{3+} + Pb$ $EMF = -0.13 - (-0.41) = +0.28\ V$

Extension

1.

In Br_2, oxidation state = 0
In HBrO, oxidation state = +1
In Br^-, oxidation state = -1

2.

Oxidation:

$Br_2 + 2H_2O \rightarrow 2HBrO + 2e^- + 2H^+$

Reduction:

$Br_2 + 2e^- \rightarrow 2Br^-$

3. Br_2 is the oxidising and reducing agent.

4. Disproportionation – bromine is both oxidised and reduced within the same reaction.

5. $+1.36\ V = +1.07\ V - E^{\ominus}_{ox}$
$E^{\ominus}_{ox} = -0.29\ V$

6. $3Br_2 + 6OH^- \rightarrow BrO_3^- + 5Br^- + 3H_2O$