Electrochemical cells misconception buster

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.

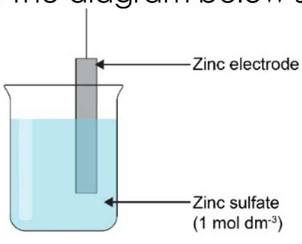
Introduction

Read each multiple choice question carefully in Activity 1 and use the diagrams and data provided to tick the correct answer. There is one correct answer per question.

After marking your multiple choice questions, complete the suggested follow-up tasks to target areas for improvement and increase your understanding of certain topics. The questions cover setting up electrochemical cells, redox equations and calculations. If you answered a question correctly but you were uncertain or guessed, make sure you complete the relevant follow-up task.

Activity 1: multiple choice questions

1. The diagram below shows:



A	A half-cell.	☐
B	An electrochemical cell.	☐
C	A battery.	☐
D	A salt bridge.	☐

2. Which concentration of sulfuric acid would you use for a standard hydrogen electrode?

- **A** $1.0\ mol\ dm^{-3}$
- **B** $2.0\ mol\ dm^{-3}$
- **C** $0.5\ mol\ dm^{-3}$
- **D** $0.25\ mol\ dm^{-3}$

3. You make an iron(II)/iron(III) half-cell. Which answer shows the IUPAC conventional representation of this half-cell when it is acting as the anode?

- **A** $Fe^{2+}(aq)\ |\ Fe^{3+}(aq)\ |\ Pt(s)$
- **B** $Fe^{2}(aq),\ Fe^{3+}(aq)\ |\ Fe(s)$
- **C** $Fe^{3+}(aq),\ Fe^{2+}(aq)\ |\ Pt(s)$
- **D** $Fe^{2+}(aq),\ Fe^{3+}(aq)\ |\ Pt(s)$

4. Which is the correct IUPAC convention for representing electrochemical cells?

- **A** Reduced form at edges, more negative cell on the left-hand side.
- **B** Reduced form at edges, more positive cell on the left-hand side.
- **C** Oxidised form at edges, more negative cell on the left-hand side.
- **D** Oxidised form at edges, more positive cell on the left-hand side.

5. Which are the correct conditions for the standard hydrogen electrode?

- **A** $298^{\circ}C$, 100 kPa, $1\ mol\ dm^{-3}$ acid solution, graphite electrode.
- **B** 298 K, 100 kPa, $1\ mol\ dm^{-3}$ H^+ solution, graphite electrode.
- **C** 298 K, 100 kPa, $1\ mol\ dm^{-3}$ H^+ solution, platinum electrode.
- **D** 298 K, 100 kPa, $1\ mol\ dm^{-3}$ acid solution, platinum electrode.

6. Why is a salt bridge used in an electrochemical cell?

- **A** It completes the cell.
- **B** It allows inert ions to transfer between half-cells to maintain electrical neutrality.
- **C** It allows electrons to transfer between half-cells to maintain electrical neutrality.
- **D** It allows current to flow between half-cells to maintain charge neutrality.

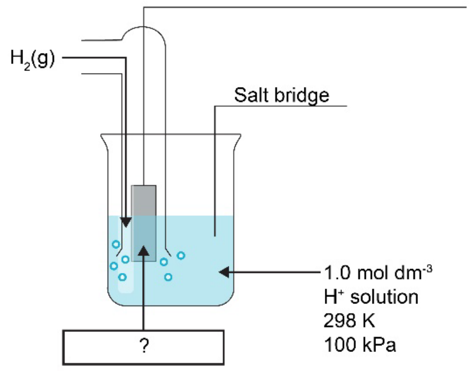
7. Which of these salts would be appropriate for producing a salt bridge?

- **A** $NaCl$
- **B** KCl
- **C** KNO_3
- **D** NH_4Cl

8. According to IUPAC convention, how is a salt bridge represented in an electrochemical cell?

- **A** ||
- **B** |
- **C** ,
- **D** /

9. The diagram represents a standard hydrogen electrode. What should the missing label read?



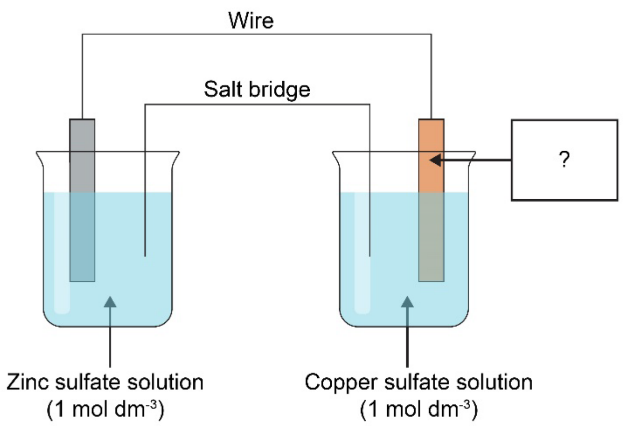
A $H_2(s)$

B $H^+(aq)$

C $Pt(s)$

D $C(s)$

10. The diagram below shows a zinc/copper electrochemical cell. What should the missing label read?



A Zinc electrode

B Platinum electrode

C Graphite electrode

D Copper electrode

11. According to IUPAC convention, which half-cell should go on the left-hand side of a cell diagram?

A The cell with the positive $E^{\ominus}$ value.

B The cell with the negative $E^{\ominus}$ value.

C The cell with the **more** positive $E^{\ominus}$ value.

D The cell with the **more** negative $E^{\ominus}$ value.

12. An iron(II)/iron(III) and zinc electrochemical cell is set up. Electrons will flow:

A From the zinc half-cell to the iron half-cell; Zn is oxidised and Fe^{3+} is reduced.

B From the zinc half-cell to the iron half-cell; Zn^{2+} is oxidised and Fe^{2+} is reduced.

C From the zinc half-cell to the iron half-cell; Fe^{2+} is oxidised and Zn^{2+} is reduced.

D From the iron half-cell to the zinc half-cell; Fe^{2+} is oxidised and Zn^{2+} is reduced.

13. Which of the standard notations below is correct according to IUPAC convention, given that $E^{\ominus}(Fe^{3+}/Fe^{2+}) = 0.77$ V and $E^{\ominus}(Zn^{2+}/Zn) = -0.76$ V?

A $Zn(s) \mid Zn^{2+}(aq) \parallel Fe^{3+}, Fe^{2+}(aq) \mid Pt(s)$

B $Zn^{2+}(aq) \mid Zn(s) \parallel Fe^{3+}, Fe^{2+}(aq) \mid Pt(s)$

C $Zn(s) \mid Zn^{2+}(aq) \parallel Fe^{3+}(aq) \mid Fe^{2+}(aq)$

D $Zn(s) \mid Zn^{2+}(aq) \parallel Fe^{2+}(aq) \mid Fe^{3+}(aq)$

A chlorine/chloride half-cell is joined to a bromine/bromide half-cell.

$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$ $E^\ominus = +1.36$ V $Br_2 + 2e^- \rightleftharpoons 2Br^-$ $E^\ominus = +1.09$ V

Use the information provided to answer questions 14–18.

14. Identify the oxidising agent.

A Cl^-

B Br_2

C Cl_2

D Br^-

15. Which species is oxidised?

A Br_2

B Br^-

C Cl_2

D Cl^-

16. Which ionic equation represents the feasible direction of the reaction?

A $Cl_2 + 2Br^- \rightleftharpoons 2Cl^- + Br_2$

B $2Cl^- + Br_2 \rightleftharpoons Cl_2 + 2Br^-$

C $Cl_2 + Br_2 \rightleftharpoons 2Cl^- + 2Br^-$

D $2Cl^- + 2Br^- \rightleftharpoons Cl_2 + Br_2$

17. Which is the positive electrode?

A Cl_2/Cl^-, as electrons flow towards it.

B Br_2/Br^-, as electrons flow towards it.

C Cl_2/Cl^-, as electrons flow away from it.

D Br_2/Br^-, as electrons flow away from it.

18. Calculate the EMF of the cell.

$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$ $E^\ominus = +1.36$ V $Br_2 + 2e^- \rightleftharpoons 2Br^-$ $E^\ominus = +1.09$ V

A +2.45 V

B +0.27 V

C -0.27 V

D -2.45 V

Reduction	$E^\ominus$ / V
$Ni^{2+} + 2e^- \rightleftharpoons Ni$	-0.25
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77
$Mg^{2+} + 2e^- \rightleftharpoons Mg$	-2.33

19. Given the $E^\ominus$ values in the table, which species is the strongest reducing agent?

A Ni^{2+}

B Cu

C Fe^{3+}

D Mg

20. Given the $E^\ominus$ values in the table, which of the following reactions are feasible?

i. $Cu + Ni^{2+} \rightleftharpoons Cu^{2+} + Ni$
ii. $Mg + 2Fe^{3+} \rightleftharpoons Mg^{2+} + 2Fe^{2+}$
iii. $Ni^{2+} + 2Fe^{2+} \rightleftharpoons Ni + 2Fe^{3+}$
iv. $Cu^{2+} + Ni \rightleftharpoons Ni^{2+} + Cu$

A i, ii, iii, and iv

B i and iii only

C ii and iv only

D None of the reactions are feasible.

Activity 2: follow-up tasks

Your teacher will give you the answers and explanations to the multiple choice questions. Using your responses to Activity 1, you will complete the suggested follow-up tasks to target areas for improvement and increase your understanding of certain topics.

Area for improvement	Follow-up task
Setting up electrochemical cells You should be able to: • Explain how to construct an electrochemical cell. • Describe the standard hydrogen electrode and explain when it is used. • Write and apply the conventional representation of an electrochemical cell and the half-equation for electrode reactions.	**(a) Constructing simple cells** A student wants to measure the electrode potential between a zinc half-cell and a silver half-cell. Describe how the student could carry this out. You should use a labelled diagram in your answer and include relevant reagents, conditions and all components of the cell. $E^\ominus(Zn^{2+}/Zn) = -0.76$ V and $E^\ominus(Ag^{+}/Ag) = +0.80$ V
	(b) Standard hydrogen electrode We measure standard electrode potentials using the standard hydrogen electrode (SHE). By definition, the SHE has an E_{cell} of 0.00 V. We measure the standard electrode potential for the reduction of iron(III) ions into iron(II) ions by connecting a suitable half-cell to a SHE: $E^\ominus(Fe^{3+}/Fe^{2+}) = 0.77$ V Draw clearly labelled diagrams to show the components and reagents, including their concentrations, in this electrochemical cell. You should include: • A diagram of the standard hydrogen electrode. • The relative positions of each half-cell in the connected electrochemical cell. • All relevant reagents, conditions and components of each half-cell.
	(c) IUPAC convention The IUPAC convention for the zinc/silver cell is given below. $E^\ominus(Zn^{2+}/Zn) = -0.76$ V and $E^\ominus(Ag^{+}/Ag) = +0.80$ V $Zn(s) \mid Zn^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$

	When drawing conventional representations of electrochemical cells: • Which cell goes on the left? • Which cells require platinum electrodes? • Why are platinum electrodes used as electrodes, rather than graphite electrodes? • Which side should the standard hydrogen electrode be placed on? • Which side of the conventional representation shows the cathode? • Which direction do electrons flow in an electrochemical cell? • What are the key components of a shortened IUPAC conventional representation?
Redox You should be able to: • Use $E^\ominus$ values to predict the feasible direction of a reaction. • Write half-equations for the reactions occurring in electrochemical cells and combine them to give the feasible direction. • Identify the anode and cathode in an electrochemical cell. • Identify oxidising and reducing agents in reactions.	**(d) Feasible direction of reactions** Step 1: identify the half-reactions. The half-reaction at the anode involves oxidation. The half-reaction at the cathode involves reduction. Step 2: assign electrode potentials to each half-cell. Step 3: calculate the EMF of the cell. $E_{cell} = E^\ominus_{red} - E^\ominus_{ox}$ Step 4: evaluate the cell potential. If E_{cell} is positive, the reaction is feasible and the electrons flow from the anode to the cathode. If E_{cell} is negative, the reaction is not feasible in these conditions. Go through the steps for the chlorine/chloride and bromine/bromide electrochemical cell provided for questions 14–18 and explain the feasible direction of the reaction. In your answer, you should: • Give the overall ionic equation for the reaction. • Identify the direction of electron flow. • Identify the cathode and anode. • Explain why the reaction is feasible under these conditions.
	(e) Writing and combining half-equations For the reactions in Q20: • Use relevant half-equations and $E^\ominus$ values to show how the ionic equations for the two feasible reactions were formed. • Use the half-equations and $E^\ominus$ values to correct the non-feasible cells. Give the appropriate direction for each half-cell and construct the full ionic equation for a feasible reaction between these half cells.

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(f) Identifying oxidation, reduction, oxidising/reducing agents, anode, cathode and direction of electron flow

Remember: AN OX RED CAT – oxidation occurs at the anode; reduction occurs at the cathode.

v. Define:
- Oxidation.
- Reducing agent.
- Cathode.

vi. For the two feasible equations in multiple choice question 20:
- Identify the oxidising and reducing agents.
- State which half-cells are the anodes and cathodes.

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Calculations

You should be able to:
- Use values to predict the direction of simple redox reactions.
- Calculate the EMF of a cell.

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(g) Calculating EMF (E_{cell})

$E_{cell} = E^{\ominus}_{red} - E^{\ominus}_{ox}$

Calculate the EMF of the electrochemical cell:

$Mg^{2+}(aq) + 2e^{-} \rightleftharpoons Mg(s)$ $E^{\ominus} = -2.38\ V$

$Ag^{+}(aq) + e^{-} \rightleftharpoons Ag(s)$ $E^{\ominus} = +0.80\ V$

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(h) Using $E^{\ominus}$ values to predict feasibility

You connect the following half-cells. Calculate the EMF for the feasible direction of this reaction.

$Cr^{3+}(aq) + e^{-} \rightleftharpoons Cr^{2+}(aq)$ $E^{\ominus} = -0.41\ V$

$Pb^{2+}(aq) + 2e^{-} \rightleftharpoons Pb(s)$ $E^{\ominus} = -0.13\ V$

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Extension

Consider the reaction of bromine in water under acidic conditions:

$2Br_2 + 2H_2O \rightarrow 2HBrO + 2Br^- + 2H^+ \quad E^\ominus = +1.36\ V$

1. Identify the oxidation states of bromine in each species.
2. Give the half-equations for the oxidation and reduction processes.
3. Identify the oxidising and reducing agent in the reaction.
4. Name this type of reaction.
5. The standard electrode potential for the reduction of bromine to bromide ions is $+1.07\ V$. Calculate the standard electrode potential for the oxidation process.
6. Under alkaline conditions, the oxidation half-equation is:

   $2Br_2 + 6OH^- \rightarrow BrO_3^- + 3Br^- + 3H_2O + 2e^-$

   Write the overall equation for the reaction.