4 Which sequence shows the ions in order of increasing strength as a reducing agent? Refer to your Data Booklet.

☒ **A** $V^{2+} < Fe^{2+} < Cr^{2+}$

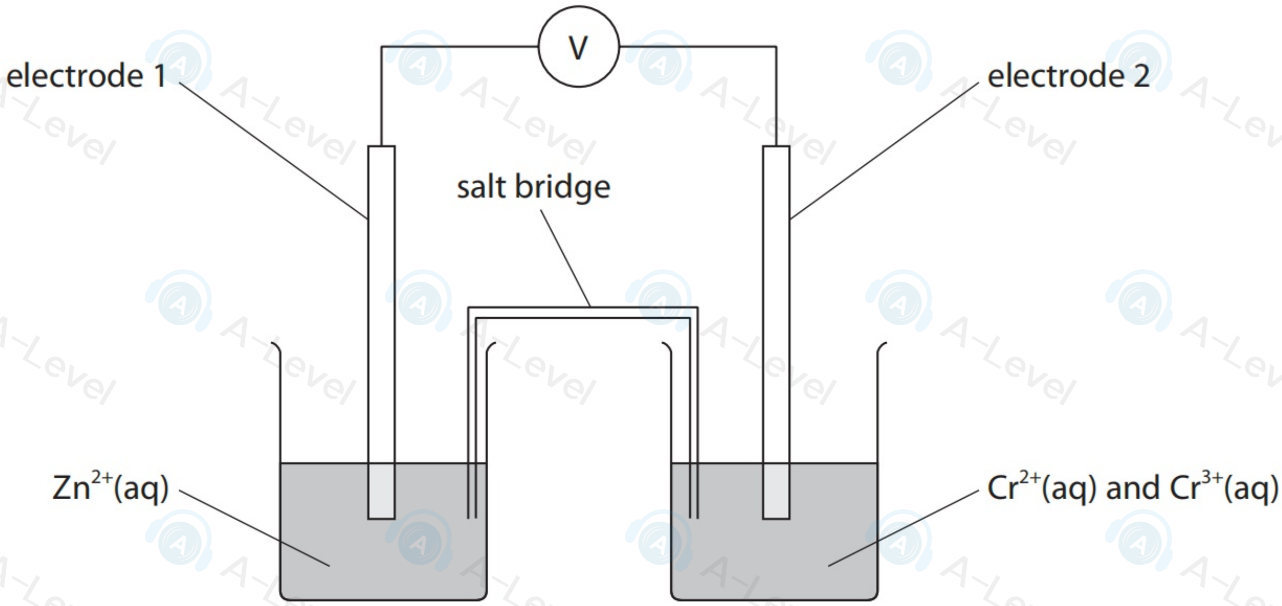
☒ **B** $Cr^{2+} < Fe^{2+} < V^{2+}$

☒ **C** $Cr^{2+} < V^{2+} < Fe^{2+}$

☒ **D** $Fe^{2+} < V^{2+} < Cr^{2+}$

1 The apparatus can be used to measure $E^{\ominus}_{cell}$ for the reaction shown.

$$2Cr^{3+}(aq) + Zn(s) \rightarrow 2Cr^{2+}(aq) + Zn^{2+}(aq)$$



(a) Which electrodes are used for this cell?

(1)

		electrode 1	electrode 2
☒	**A**	platinum	platinum
☒	**B**	platinum	chromium
☒	**C**	zinc	chromium
☒	**D**	zinc	platinum

(b) A student wishes to measure the standard cell potential, $E^{\ominus}_{cell}$, of this cell. The right-hand cell requires Cr^{3+} and Cr^{2+} ions.

What mass of $Cr_2(SO_4)_3 \cdot 18H_2O$ must be dissolved in 1.00 dm³ of deionised water to give the concentration of Cr^{3+} ions required to measure this $E^{\ominus}_{cell}$?

(1)

☒ **A** 52.0 g

☒ **B** 196 g

☒ **C** 358 g

☒ **D** 716 g

(c) What can be deduced from the fact that, for this reaction, $E^{\ominus}_{cell}$ is positive?

(1)

☒ **A** ΔS_{total} and $\ln K$ are positive

☒ **B** ΔS_{total} and $\ln K$ are negative

10 The emf, $E^{\ominus}_{cell}$, of a cell is +0.57 V.

The numerical values of the standard electrode potentials of the two half-cells joined in this cell are 0.17 V and 0.40 V.

What are the signs of the standard electrode potentials of the right-hand half-cell and the left-hand half-cell?

	Sign of standard electrode potential	
	in left-hand half-cell	in right-hand half-cell
☒ **A**	negative	negative
☒ **B**	negative	positive
☒ **C**	positive	negative
☒ **D**	positive	positive

2 Some standard electrode potentials are shown.

Right-hand electrode system	$E^{\ominus}$ / V
$Mg^{2+} + 2e^{-} \rightleftharpoons Mg$	−2.37
$Ce^{3+} + 3e^{-} \rightleftharpoons Ce$	−2.33
$Mn^{2+} + 2e^{-} \rightleftharpoons Mn$	−1.19
$Mn^{3+} + e^{-} \rightleftharpoons Mn^{2+}$	+1.49
$Ce^{4+} + e^{-} \rightleftharpoons Ce^{3+}$	+1.70

Which reaction is thermodynamically feasible?

☒ **A** $2Ce + 3Mg^{2+} \rightarrow 2Ce^{3+} + 3Mg$

☒ **B** $4Ce^{3+} \rightarrow Ce + 3Ce^{4+}$

☒ **C** $3Mn^{2+} \rightarrow Mn + 2Mn^{3+}$

☒ **D** $Mg + 2Ce^{4+} \rightarrow Mg^{2+} + 2Ce^{3+}$

4 For the reaction shown, $E^{\ominus}_{cell} = +0.89V$.

$$4V^{3+}(aq) + O_2(g) + 2H_2O(l) \rightarrow 4VO^{2+}(aq) + 4H^+(aq)$$

Which statement can be deduced from this information only?

☒ **A** the reaction will **not** occur under any conditions

☒ **B** the reactants are kinetically stable with respect to the products

☒ **C** the reactants are thermodynamically unstable with respect to the products

☒ **D** an aqueous solution of V^{3+} will oxidise rapidly on standing in air

11 An electrochemical cell is set up:

- a half-cell is made from a piece of zinc and a solution of zinc chloride, $ZnCl_2$
- a second half-cell is made from a piece of metal **G** and a solution of its chloride, **G**Cl_2
- the two half-cells are connected and a current allowed to pass for some time.

The zinc electrode increased in mass by 1.635 g.

The electrode of metal **G** decreased in mass by 0.6075 g.

What is metal **G**?

☒ **A** copper

☒ **B** iron

☒ **C** magnesium

☒ **D** manganese

3 In a chemical reaction, $E^{\ominus}_{cell}$ is **directly** proportional to

☒ **A** K_c

☒ **B** $\Delta_r H^{\ominus}$

☒ **C** $\Delta S^{\ominus}_{system}$

☒ **D** $\Delta S^{\ominus}_{total}$

8 The cell diagram for an electrochemical cell is shown.

$$Pt(s) \mid Fe^{2+}(aq), Fe^{3+}(aq) \parallel Br_2(aq), 2Br^-(aq) \mid Pt(s) \qquad E^{\ominus}_{cell} = +0.32\,V$$

What are the half-equations for the reactions that occur in the half-cells?

		Half-equation in left-hand half-cell	Half-equation in right-hand half-cell
☒	**A**	$Fe^{2+}(aq) \rightleftharpoons Fe^{3+}(aq) + e^-$	$2Br^-(aq) \rightleftharpoons Br_2(aq) + 2e^-$
☒	**B**	$Fe^{2+}(aq) \rightleftharpoons Fe^{3+}(aq) + e^-$	$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq)$
☒	**C**	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	$2Br^-(aq) \rightleftharpoons Br_2(aq) + 2e^-$
☒	**D**	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq)$

1 The equation for a redox reaction is shown.

$$2V^{2+}(aq) + Cu^{2+}(aq) \rightarrow 2V^{3+}(aq) + Cu(s)$$

What is the cell diagram for this reaction?

☒ **A** $V(s) \mid V^{3+}(aq), V^{2+}(aq) \parallel Cu(s) \mid Cu^{2+}(aq)$

☒ **B** $V(s) \mid V^{2+}(aq), V^{3+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

☒ **C** $Pt(s) \mid V^{3+}(aq), V^{2+}(aq) \parallel Cu(s) \mid Cu^{2+}(aq)$

☒ **D** $Pt(s) \mid V^{2+}(aq), V^{3+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

8 Standard electrode potentials can also be given an alternative name.
The electrochemical series lists standard electrode potentials in order.

Which of these is correct?

		Alternative name for standard electrode potential	Order of standard electrode potentials in the electrochemical series
☒	**A**	standard reduction potential	most negative to most positive
☒	**B**	standard reduction potential	most positive to most negative
☒	**C**	standard cell potential	most negative to most positive
☒	**D**	standard cell potential	most positive to most negative

7 $E^{\ominus}_{cell}$ is directly proportional to

☒ **A** $\Delta_r H$ and $\ln K$

☒ **B** $\Delta_r H$ and $\ln RT$

☒ **C** ΔS_{total} and $\ln K$

☒ **D** ΔS_{total} and $\ln RT$

21 This question is about mercury, Hg, and its compounds.

Mercury is a liquid element in the same group of the Periodic Table as zinc.

The electronic configuration of mercury is $[Xe]4f^{14}5d^{10}6s^{2}$.

(a) Mercury forms compounds in either the +1 or +2 oxidation states.

Explain why mercury is classified as a d-block element but is **not** a transition element.

(3)

(b) Mercury reacts with nitric acid to form an aqueous solution of $Hg(NO_3)_2$ and nitrogen monoxide gas.

The **unbalanced** equation is shown.

$$Hg(l) + HNO_3(aq) \rightarrow Hg(NO_3)_2(aq) + NO(g) + H_2O(l)$$

(i) Explain, using oxidation numbers, why this is a redox reaction.

(2)

....

....

....

....

....

(ii) Deduce the **ionic** half-equations for this reaction.
State symbols are not required.

(2)

(iii) Complete the equation for this reaction by adding the stoichiometric coefficients.

(1)

$$\ldots\ldots Hg(l) + \ldots\ldots HNO_3(aq) \rightarrow \ldots\ldots Hg(NO_3)_2(aq) + \ldots\ldots NO(g) + \ldots\ldots H_2O(l)$$

(c) Mercury(II) fulminate, $Hg(CNO)_2$, is an explosive.

(i) It is produced in the reaction between $Hg(NO_3)_2$ and ethanol. The other products of the reaction are ethanal and water.

Write the equation for the reaction of one mole of $Hg(NO_3)_2$ with ethanol to form mercury(II) fulminate.
State symbols are not required.

(2)

(ii) $Hg(CNO)_2$ decomposes as shown.

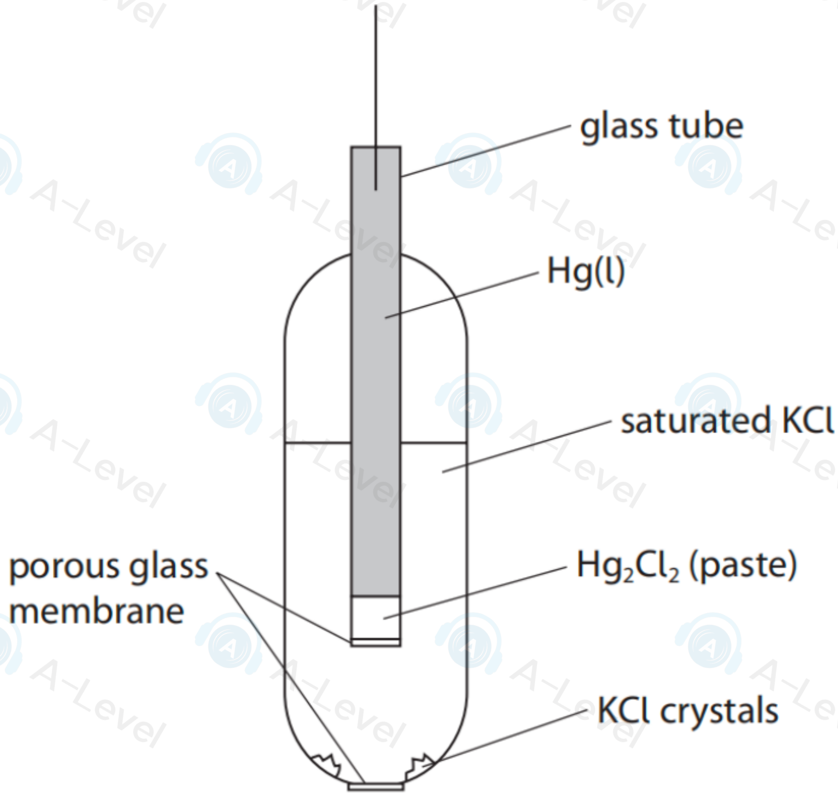
$$3Hg(CNO)_2(s) \rightarrow Hg(CN)_2(s) + 2Hg(s) + 2CO_2(g) + 2CO(g) + 2N_2(g)$$

Calculate the **total** volume, in cm^3, of gas produced when 1.00 g of $Hg(CNO)_2$ decomposes at room temperature and pressure.

(3)

(d) Mercury(I) chloride, Hg_2Cl_2, is also known as calomel.

A saturated calomel electrode may be used as an alternative to the standard hydrogen electrode.

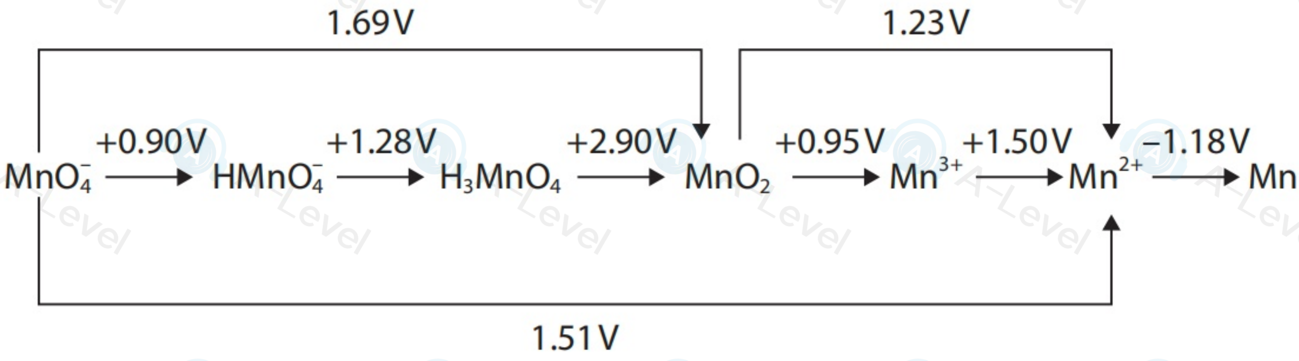


The half-equation for the calomel electrode is

16 A Latimer diagram for a chemical element is a summary of the standard electrode potential data for that element.

In a Latimer diagram, the form of the element with the highest oxidation state is on the left, with successively lower oxidation states to the right.

A Latimer diagram for manganese at pH = 0 is shown.



The diagram shows that the standard electrode potential for the reduction of MnO_4^- to MnO_2, in acidic conditions, is +1.69V.

$$MnO_4^- + 4H^+ + 3e^- \rightleftharpoons MnO_2 + 2H_2O \qquad E^\ominus = +1.69\,V$$

(a) (i) Justify the assignment of the oxidation state of +5 to manganese in H_3MnO_4 using oxidation numbers.

(1)

(ii) Write an equation for the reaction of H_3MnO_4 in acidic solution to give ions containing manganese(VI) and manganese(IV).
Use the Latimer diagram to obtain the formulae of the ions produced.
State symbols are not required.

(2)

(iii) Deduce whether or not this disproportionation reaction is thermodynamically feasible by calculating $E^\ominus$ for the reaction.

(2)

(b) Before use in titration experiments, potassium manganate(VII) solutions must be standardised. One method uses ethanedioate ions to find the exact concentration of the manganate(VII) ions.

250.0 cm^3 of a standard solution contained 1.915 g of sodium ethanedioate, $Na_2C_2O_4$.

A potassium manganate(VII) solution of approximately 0.02 mol dm^{-3} was standardised using this solution.
Excess sulfuric acid was added to 25.0 cm^3 portions of the potassium manganate(VII) solution which were titrated with the sodium ethanedioate solution.

The mean titre was 22.95 cm^3.

The relevant ionic half-equations are shown.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$$

(i) State the colour change at the end-point of the titration.

(1)

....

....

(ii) Calculate the accurate concentration of the potassium manganate(VII), in mol dm^{-3}, giving your answer to an appropriate number of significant figures.

(4)

(iii) A second titration carried out without the addition of sulfuric acid resulted in the formation of a brown suspension.
Explain how the value of the mean titre would be affected, if at all, by the reaction that forms this suspension.
Use the Data Booklet as a source of information.
There is no need to calculate E_{cell} values.

(3)

18 This question is about electrochemical cells and redox reactions.

(a) Draw a labelled diagram of the apparatus you would use to measure the standard electrode potential of a $Cu^{2+}(aq)|Cu(s)$ electrode with a standard hydrogen electrode.
Include essential conditions.

(5)

(b) The standard electrode potentials for two half-cells are shown.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l) \qquad E^\ominus = +1.33\,V$$

$$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq) \qquad E^\ominus = +1.36\,V$$

(i) Explain, in terms of electrode potentials, why acidified dichromate(VI) ions react with **concentrated** hydrochloric acid to form chlorine, even though $E^\ominus_{cell} = -0.03\,V$.

(3)

(ii) Write the cell diagram, using the conventional representation of half-cells, for the reaction to produce chlorine.

(2)

(c) When $25.0\,cm^3$ of a $0.100\,mol\,dm^{-3}$ solution of X_2O_5 reacts with a reducing agent, X is reduced to a lower oxidation state.

To oxidise X back to its original oxidation state required $50.0\,cm^3$ of $0.0200\,mol\,dm^{-3}$ acidified potassium manganate(VII) solution.

The half-equation for acidified manganate(VII) is

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

Calculate the oxidation state of X after it has been reduced.

(3)