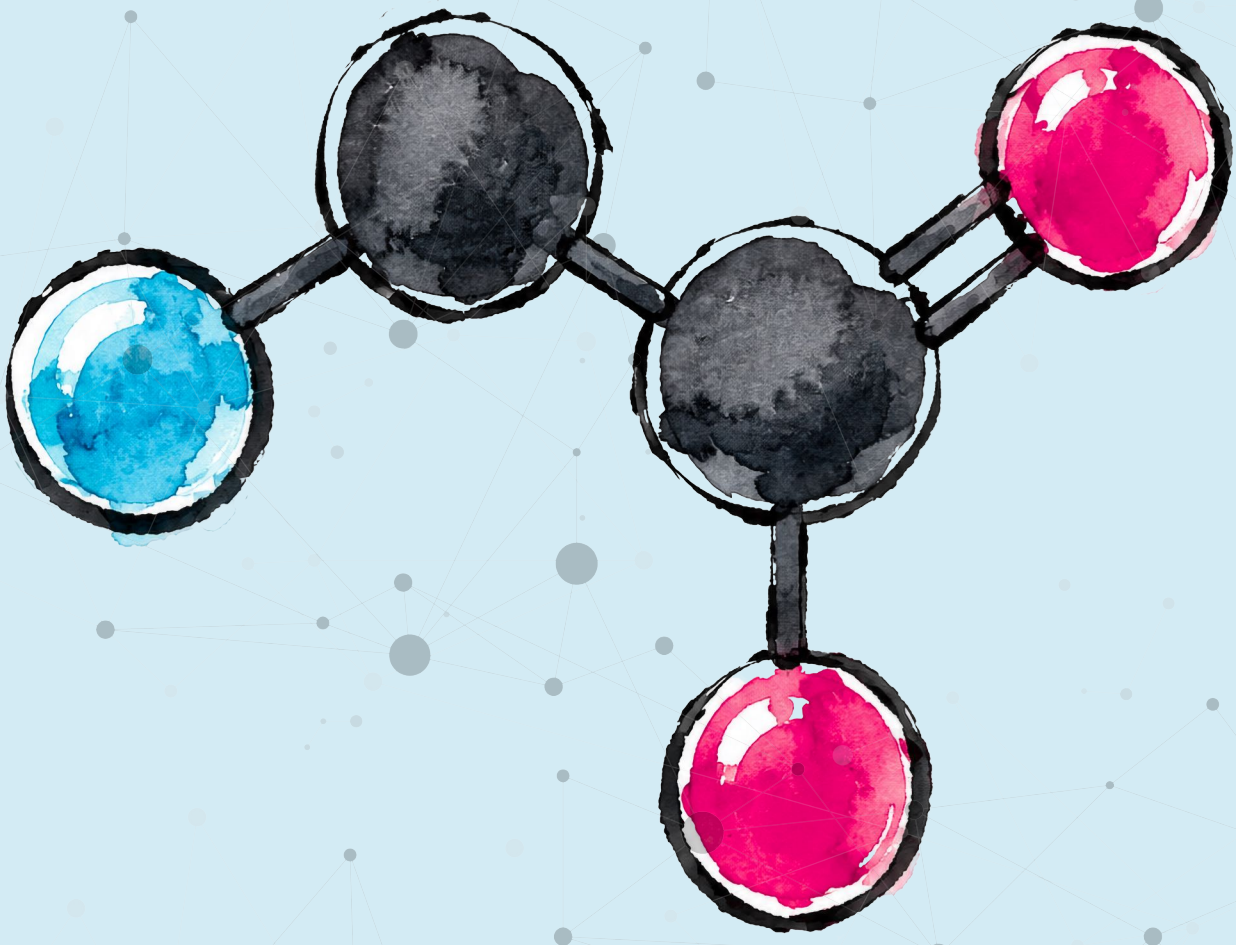
edexcel iAL

A2 CHEMISTRY



Topical Questions

Unit 5

Edexcel IAL Chemistry

Unit 5 (WCH15/01) • Questions and Answers by Topic

Past-paper questions classified by topic

Edexcel IAL - Solved Past Paper

October 2020 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q14

hard medium easy

14

Which · [1 mark]

Which species contains an element with the same oxidation number as bromine in KBrO3?

Answer

- [x] **A** N2O5
- [ ] **B** KMnO4
- [ ] **C** K2FeO4
- [ ] **D** Na2SO3

The logic

✓ In KBrO3: K is +1 and each O is -2 (total -6), so Br = +6 - 1 = +5.

In N2O5: each O is -2 (total -10), shared over 2 N, so each N = +5 - matching bromine.

Why the others fail

A

B in KMnO4, Mn = +7, not +5.

C in K2FeO4, Fe = +6.

D in Na2SO3, S = +4.

Where marks are lost

Rushing the arithmetic. Assign O = -2 and the Group 1 metal = +1, then solve for the target element in each option.

Examiner's insight

A quick oxidation-number balance on each species - only N in N2O5 comes out at +5.

Revise from your notes

- Grignard reagents 10.4
- Identifying structures from spectra 10.10
- Oxidation number and redox 6.1

Edexcel IAL - Solved Past Paper

October 2020 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q15

hard medium easy

15

Which · [1 mark]

Which is a redox reaction? (A) $Cr_2O_7^{2-} + 2C \longrightarrow Cr_2O_3 + CO_3^{2-} + CO$ (B) $CrO_2Cl_2 + 4OH^- \longrightarrow CrO_4^{2-} + 2Cl^- + 2H_2O$ (C) $Cr_2O_7^{2-} + 2HCl \longrightarrow 2CrO_3Cl^- + H_2O$ (D) $2CrO_4^{2-} + 2H^+ \longrightarrow Cr_2O_7^{2-} + H_2O$

Answer

- [x] **A** $Cr_2O_7^{2-} + 2C \longrightarrow Cr_2O_3 + CO_3^{2-} + CO$
- [ ] **B** $CrO_2Cl_2 + 4OH^- \longrightarrow CrO_4^{2-} + 2Cl^- + 2H_2O$
- [ ] **C** $Cr_2O_7^{2-} + 2HCl \longrightarrow 2CrO_3Cl^- + H_2O$
- [ ] **D** $2CrO_4^{2-} + 2H^+ \longrightarrow Cr_2O_7^{2-} + H_2O$

The logic

✓ A redox reaction needs at least one element to change oxidation number. Track chromium and carbon in A.

In A, chromium falls from +6 ($Cr_2O_7^{2-}$) to +3 (Cr_2O_3) - reduction - while carbon rises from 0 (C) to +4 (CO_3^{2-}) and +2 (CO) - oxidation. Both change, so A is redox.

Why the others fail

B chromium stays +6 ($CrO_2Cl_2 \longrightarrow CrO_4^{2-}$) and no other atom changes - this is not redox.

C chromium stays +6 throughout - no oxidation-number change, so not redox.

D chromium stays +6 (chromate ⟶ dichromate); this is just an acid-base equilibrium, not redox.

Where marks are lost

Assuming any chromium reaction is redox. Chromate/dichromate interconversions keep Cr at +6 - check every element's oxidation number changes before calling it redox.

Examiner's insight

Only option A changes an oxidation number (Cr down, C up); the other three are rearrangements of chromium(VI).

Revise from your notes

- Grignard reagents 10.4
- Identifying structures from spectra 10.10
- Oxidation number and redox 6.1

Edexcel IAL - Solved Past Paper

October 2020 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q20

hard medium easy

20

What · [1 mark]

A solution contains 19.6 g dm^{-3} of chromium(III) sulfate, Cr2(SO4)3. What is the concentration, in mol dm^{-3}, of sulfate ions in this solution? [Molar mass Cr2(SO4)3 = 392 g mol^{-1}]

Answer

- [ ] **A** 0.05
- [ ] **B** 0.10
- [x] **C** **0.15**
- [ ] **D** 0.25

How to get there

✓ concentration of Cr2(SO4)3 = mass / M_r = 19.6 / 392 = 0.0500 mol dm^{-3}

✓ each formula unit releases 3 sulfate ions, so [SO_4^{2-}] = 3 x 0.0500 = 0.15 mol dm^{-3}

0.15 mol dm^{-3}

Why the others fail

A 0.05 is the concentration of Cr2(SO4)3 itself - it forgets that each unit gives three sulfates.

B 0.10 is the concentration of chromium(III) ions (2 x 0.05), not sulfate.

D 0.25 is the total concentration of ALL ions (2 Cr^{3+} + 3 SO_4^{2-} = 5 x 0.05).

Where marks are lost

Stopping at 0.05. The formula has THREE sulfates per unit, so the sulfate concentration is three times the salt concentration.

Examiner's insight

Read the subscript: Cr2(SO4)3 gives 3 sulfate ions per formula unit, so multiply the salt concentration by three.

Revise from your notes

- Electron configurations and variable oxidation number 7
- Ligands, complex ions and coordination number 7
- Redox titration calculations 6.9

Edexcel IAL - Solved Past Paper

October 2020 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q21

hard medium easy

21 (a)

Identify · [3 marks]

The standard electrode potential of an iron(III)/iron(II) half-cell, $Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$, is measured against a standard hydrogen electrode. Identify, by name or formula, the substances needed in the salt bridge (A), the right-hand electrode (B) and the right-hand half-cell solution (C).

Model answer

✓ A – Salt bridge: a solution of potassium nitrate, KNO3 (any soluble salt that does not react will do – KCl, NaNO3, NaCl or an ammonium salt). ✓ B – Electrode: platinum, Pt. The Fe^{3+}/Fe^{2+} system is two ions in solution, so an inert platinum electrode carries the electrons. ✓ C – Solution: contains BOTH iron(II) and iron(III) ions, Fe^{2+} and Fe^{3+} (as soluble compounds such as the chlorides, nitrates or sulfates), each at 1.00 mol dm^{-3}.

Mark scheme 3 marks - one each for A, B and C

A

✓ potassium nitrate / KNO3 in the salt bridge (1)

B

✓ platinum / Pt electrode (1)

C

✓ solution containing both Fe^{2+} and Fe^{3+} ions (1)

Accept: KCl / NaNO3 / NaCl / ammonium salts for the salt bridge; soluble iron(II) and iron(III) compounds (chlorides, nitrates, sulfates) for C

Ignore: any conditions or concentrations quoted; mention of acid in C

Do not accept: just 'iron' for the electrode - the half-cell is a solution of two ions, needing an INERT electrode

How to reach full marks

Three separate marks. The one most often missed is C: the right-hand solution must contain BOTH oxidation states of iron, because the half-cell is the Fe^{3+}/Fe^{2+} equilibrium, not iron metal.

Where marks are lost

Writing 'iron' as the electrode. There is no iron metal in an Fe^{3+}/Fe^{2+} half-cell - it needs an inert Pt electrode dipping into a solution of both ions.

Examiner's insight

Most candidates named the salt bridge and platinum; fewer realised the right-hand solution must hold both iron(II) and iron(III).

21 (b)

Write · [3 marks]

The cell diagram is Pt | [Mn^{2+} + 4H2O], [$MnO4^-$ + $8H^+$] || [$BiO3^-$ + $6H^+$], [Bi^{3+} + 3H2O] | Pt, with E_{cell} = +0.09 V. Write the half-equation for each half-cell and hence the overall ionic

equation. State symbols are not required.

The two half-equations, then the overall equation

✓ Bismuthate half-cell (reduction): $BiO_3^- + 6H^+ + 2e^- \longrightarrow Bi^{3+} + 3H_2O$

✓ Manganese half-cell (oxidation, so written in reverse): $Mn^{2+} + 4H_2O \longrightarrow MnO_4^- + 8H^+ + 5e^-$

E_{cell} = +0.09 V (positive) means bismuthate is reduced and Mn^{2+} is oxidised. Scale to cancel electrons: bismuthate x 5 (10 e^-), manganese x 2 (10 e^-), then cancel the spectator H^+ and H_2O.

✓ Overall: $2Mn^{2+} + 5BiO_3^- + 14H^+ \longrightarrow 2MnO_4^- + 5Bi^{3+} + 7H_2O$

$$2Mn^{2+} + 5BiO_3^- + 14H^+ \longrightarrow 2MnO_4^- + 5Bi^{3+} + 7H_2O$$

Mark scheme 3 marks - two half-equations and the overall equation

- ✓ half-equation for bismuthate ions (allow reversed, allow the equilibrium sign) (1)
- ✓ half-equation for manganate(VII) ions (allow reversed, allow -5e^- on the left) (1)
- ✓ overall equation, written in the direction shown, electrons/H^+/H_2O fully cancelled (1)

Accept: half-equations written in reverse; the equilibrium sign in place of the arrow; multiples of the overall equation

Do not accept: uncancelled electrons, H^+ or H_2O left in the overall equation

Ignore: state symbols, even if incorrect

How to reach full marks

The overall-equation mark needs the electrons balanced (x5 and x2 to reach 10 each) AND the leftover H^+ and H_2O cancelled to one side. The direction is fixed by the positive E_{cell}: manganese is oxidised, bismuthate reduced.

Where marks are lost

Omitting the electrons or mis-counting them in the half-equations, then failing to cancel the 16 H^+ down to 14 and 8 H_2O down to 7 when combining.

Examiner's insight

This was challenging: many dropped the electrons or could not turn the two half-equations and the E_{cell} sign into a single balanced overall equation.

21 (c)

Calculate · [2 marks]

For $Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$, E-standard = -0.74 V. Using E = E-standard + (RT / (96500 x n)) x ln[Cr^{3+}], calculate the electrode potential at 298 K when [Cr^{3+}] = 0.0100 mol dm^{-3}. (R = 8.31 J K^{-1} mol^{-1}, n = 3.)

Working

✓ substitute: E = -0.74 + (8.31 x 298) / (96500 x 3) x ln(0.0100)

the front factor = 2476.4 / 289500 = 0.008554; ln(0.0100) = -4.6052

E = -0.74 + 0.008554 x (-4.6052) = -0.74 - 0.0394

✓ E = -0.779 V

E = -0.779 V

Mark scheme 2 marks - substitution and answer

- ✓ correct substitution of values into the formula (n = 3, $[Cr^{3+}]$ = 0.0100) (1)
- ✓ E = -0.779 V (accept -0.78 to -0.7794) (1)

Accept: a correct answer with no working scores both marks; transferred error on wrong numbers put into the correct formula, e.g. $[Cr^{3+}]$ = 0.100 gives -0.76 V

Ignore: units, and significant figures except 1 sf

Do not accept: no transferred error if the FORMULA itself is altered

How to reach full marks

Two marks: a clean substitution then the value. The two numbers most often slipped are n (it is 3, from the half-equation) and $[Cr^{3+}]$ (0.0100, not 0.100) - both are given in the question.

Where marks are lost

Using n = 1, or reading the concentration as 0.100. Both were printed in the question, so these are avoidable slips.

Examiner's insight

Most substituted correctly and reached the right value; the lost marks came from careless use of n or the concentration.

Revise from your notes

- Standard electrode potential and the SHE 6.2
- Cell diagrams and combining electrode potentials 6.2
- Oxidation numbers and half-equations 6.1

Edexcel IAL - Solved Past Paper

January 2021 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q1

hard medium easy

1

How · [1 mark]

When an alkene is added to potassium manganate(VII), the purple solution turns colourless. In terms of electron transfer and oxidation number, how does the manganese change?

Answer

- [ ] **A** gains electrons, oxidation number increases
- [x] **B** **gains electrons, oxidation number decreases**
- [ ] **C** loses electrons, oxidation number increases
- [ ] **D** loses electrons, oxidation number decreases

The logic

✓ The purple colour is the manganate(VII) ion MnO_4^-, in which manganese is +7. It fades to the almost colourless Mn^{2+} ion, so manganese has gone from +7 down to +2 – the oxidation number DECREASES.

A fall in oxidation number is **reduction**, and **reduction** is a GAIN of electrons. So manganese gains electrons and its oxidation number decreases.

Why the others fail

A the electron half is right (reduction is a gain), but a species being reduced has its oxidation number FALL, not rise.

C 'loses electrons, increases' describes oxidation - the exact opposite of what happens to manganese here.

D the oxidation number does fall, but reduction is a GAIN of electrons, so 'loses electrons' is wrong.

Where marks are lost

Pairing the gain of electrons with a rise in oxidation number. Gaining electrons is reduction, and reduction always LOWERS the oxidation number.

Examiner's insight

Link the two ideas the same way every time: reduction = gain of electrons = fall in oxidation number.

Revise from your notes

- Oxidation number and half-equations 6.1
- The standard hydrogen electrode 6.2
- Fuel cells 6.11
- Transition metals: definition and electron configurations 7

Edexcel IAL - Solved Past Paper

January 2021 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q2

hard medium easy

2

What · [1 mark]

The standard hydrogen electrode uses a platinum electrode coated in a finely divided form of the metal called platinum black. What is the purpose of this coating?

Answer

- [x] **A** **to increase the rate of the equilibrium between the hydrogen gas and the hydrogen ions**
- [ ] **B** to provide an inert protective coating for the electrode
- [ ] **C** to increase the electrical conductivity of the electrode
- [ ] **D** to ensure that the conditions remain standard

The logic

✓ Platinum black is platinum deposited in a very finely divided form, which gives it a huge surface area.

That large surface acts as a catalyst, letting the $H_2(g)$ / $H^+(aq)$ equilibrium at the electrode be reached quickly so the electrode gives a steady, reliable potential.

Why the others fail

A

B platinum black is chemically identical to shiny platinum and is already inert - it is not there as protection.

C it is the same metal, so it does not change the electrical conductivity.

D the coating does not alter temperature, pressure or concentration, so it is not what keeps the conditions standard.

Where marks are lost

Thinking the coating is protective or improves conductivity. Its job is the large surface area, which speeds up (catalyses) the electrode equilibrium.

Examiner's insight

Finely divided platinum means a large surface area, which catalyses the establishment of the hydrogen electrode equilibrium.

Revise from your notes

- Oxidation number and half-equations 6.1
- The standard hydrogen electrode 6.2
- Fuel cells 6.11
- Transition metals: definition and electron configurations 7

Edexcel IAL - Solved Past Paper

January 2021 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q3

hard medium easy

3 (a)

What · [1 mark]

An electrochemical cell is set up to measure E-cell for the reaction $Fe(s) + Sn^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Sn(s)$. What is the cell diagram for this cell?

Answer

- [ ] **A** $Fe(s) \mid Fe^{2+}(aq) \parallel Sn(s) \mid Sn^{2+}(aq)$
- [ ] **B** $Fe^{2+}(aq) \mid Fe(s) \parallel Sn^{2+}(aq) \mid Sn(s)$
- [x] **C** $Fe(s) \mid Fe^{2+}(aq) \parallel Sn^{2+}(aq) \mid Sn(s)$
- [ ] **D** $Fe^{2+}(aq) \mid Fe(s) \parallel Sn(s) \mid Sn^{2+}(aq)$

The logic

✓ Read the equation: iron is oxidised ($Fe \longrightarrow Fe^{2+}$) so it is the LEFT electrode; tin is reduced ($Sn^{2+} \longrightarrow Sn$) so it is the RIGHT electrode.

In cell notation the two OXIDISED species sit next to the central salt bridge (||). Left: $Fe(s) \mid Fe^{2+}$, right: $Sn^{2+} \mid Sn(s)$. That gives $Fe(s) \mid Fe^{2+}(aq) \parallel Sn^{2+}(aq) \mid Sn(s)$.

Why the others fail

- **A** the left side is right, but the right half-cell is written $Sn(s) \mid Sn^{2+}$ - the oxidised species Sn^{2+} must be next to the salt bridge, not on the outside.
- **B** the left half-cell is reversed ($Fe^{2+} \mid Fe$ puts the reduced form at the junction); both half-cells are the wrong way round.
- **D** both half-cells are reversed relative to the correct diagram.

Where marks are lost

Getting the order within a half-cell wrong. The species that is oxidised (Fe^{2+} on the left, Sn^{2+} on the right) always sits next to the salt bridge.

Examiner's insight

Oxidation on the left, reduction on the right, and the two oxidised species meet in the middle at the salt bridge.

3 (b)

What · [1 mark]

The standard electrode potential for the Fe / Fe^{2+} system is -0.44 V and E-cell for the reaction $Fe(s) + Sn^{2+} \longrightarrow Fe^{2+} + Sn(s)$ is +0.30 V. What is the standard electrode potential for the Sn / Sn^{2+} system?

Answer

- [ ] **A** -0.74 V

- [x] **B** -0.14 V
- [ ] **C** +0.14 V
- [ ] **D** +0.74 V

How to get there

✓ E-cell = E(right) - E(left). Iron is oxidised, so it is the left electrode: E(left) = -0.44 V. Tin is the right electrode: E(right) = $E(Sn/Sn^{2+})$.

Rearrange: $E(Sn/Sn^{2+})$ = E-cell + E(left) = (+0.30) + (-0.44) = -0.14 V.

$E(Sn / Sn^{2+})$ = -0.14 V

Why the others fail

A -0.74 V comes from SUBTRACTING E-cell from the iron value (-0.44 - 0.30) instead of adding it.
C +0.14 V is the right size but the sign has been reversed.
D +0.74 V combines both slips: subtracting E-cell (-0.74) and then reversing the sign.

Where marks are lost

Subtracting the E-cell value instead of adding it, or flipping the sign of the answer. Keep to E-cell = E(right) - E(left) and rearrange carefully.

Examiner's insight

Set out E-cell = E(right) - E(left) with iron on the left, then rearrange - the common errors were a subtraction or a sign flip.

Revise from your notes

- Cell diagrams and standard EMF 6.4
- Standard electrode potential and combining half-cells 6.2 6.6

Edexcel IAL - Solved Past Paper

June 2022 · Unit 5: Transition Metals and Organic Nitrogen Chemistry (WCH15/01) · Q17

hard medium easy

17 (a)(i)

Determine · [4 marks]

2.72 g of brass is dissolved and its copper is converted to CuSCN (M = 121.6) via $Cu^{2+} \longrightarrow Cu^{+} \longrightarrow$ CuSCN. The precipitate of CuSCN has a mass of 4.69 g. A gilding metal contains 5-11 % zinc by mass. Determine whether this brass is a gilding metal by calculating its percentage by mass of copper.

Working

- ✓ moles of CuSCN = mass / M = 4.69 / 121.6 = 0.038569 mol
- ✓ the ratio Cu : CuSCN is 1 : 1, so moles of Cu = 0.038569 mol
- ✓ mass of Cu = moles x A(Cu) = 0.038569 x 63.5 = 2.4491 g
- ✓ percentage of Cu = (2.4491 / 2.72) x 100 = 90.0 %, so zinc = 10.0 % - within 5-11 %, so it IS a gilding metal

90.0 % copper (10.0 % zinc) - it is a gilding metal

Mark scheme 4 marks

- ✓ moles of CuSCN = 0.0386 mol (1)
- ✓ moles of Cu = 0.0386 mol (1 : 1 ratio - may be subsumed into the mass step) (1)
- ✓ mass of Cu = 2.45 g (1)
- ✓ percentage of Cu = 90.0 % AND the statement that it is a gilding metal (1)

Accept: 121.5 for the M of CuSCN in M1, but not 64 for copper in M3; transferred error at each step; a percentage of zinc used to decide; ignore significant figures except 1 sf

Do not accept: transferred error for M4 giving above 100 %; a percentage outside 89-95 % that is still called a gilding metal

How to reach full marks

Moles of CuSCN, then 1:1 to moles of Cu, then mass, then percentage over 2.72 g. The fourth mark needs BOTH the percentage AND the conclusion - 10 % zinc lies in the 5-11 % range, so it is a gilding metal.

Where marks are lost

Using an atomic mass of 64 (or the wrong M for CuSCN), or calculating the percentage but forgetting to STATE the conclusion. Note the answer is the copper percentage - subtract from 100 for the zinc.

Examiner's insight

Handled well; the usual error was a wrong molar mass, and full marks needed the explicit statement that this brass is a gilding metal.

17 (a)(ii)

Explain · [3 marks]

Explain, using the data, why $HSO3^{-}$ reduces Cu^{2+} to Cu^{+} but not Cu^{+} to Cu, considering

both thermodynamic and kinetic factors. Data: Cu^{2+}/Cu^{+} = +0.15 V; $HSO4^{-}/HSO3^{-}$ = +0.17 V; Cu^{+}/Cu = +0.52 V.

The two E-cell values, then the interpretation

for $HSO3^{-}$ reducing Cu^{2+} to Cu^{+}: E-cell = $E(Cu^{2+}/Cu^{+})$ - $E(HSO4^{-}/HSO3^{-})$ = 0.15 - 0.17 = -0.02 V

for $HSO3^{-}$ reducing Cu^{+} to Cu: E-cell = $E(Cu^{+}/Cu)$ - $E(HSO4^{-}/HSO3^{-})$ = 0.52 - 0.17 = +0.35 V

- ✓ So on the standard data, Cu^{2+} should NOT be reduced to Cu^{+} (E-cell negative) while Cu^{+} SHOULD be reduced to Cu (E-cell positive) - the opposite of what is seen.
- ✓ Cu^{2+} IS reduced to Cu^{+} because the conditions are non-standard and the two E values are very close (-0.02 V), so a small shift makes it feasible.
- ✓ Cu^{+} is NOT reduced to Cu (despite the positive E-cell) because that reaction is kinetically hindered - a high activation energy makes it too slow.

E-cell = -0.02 V and +0.35 V; non-standard conditions allow the first, kinetics blocks the second

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Mark scheme 3 marks

- ✓ the E-cell data indicate Cu^{2+} should NOT be reduced to Cu^{+} AND Cu^{+} should be reduced to Cu (1)
- ✓ Cu^{2+} can be reduced to Cu^{+} because conditions are non-standard and the E values are so close (1)
- ✓ Cu^{+} is not reduced to Cu because the reaction is kinetically hindered / high activation energy / very slow (1)

Accept: E-cell = -0.02 V and E-cell = +0.35 V; addition of OH^{-} ions as an alternative reason for non-standard conditions; 'not kinetically favoured'

How to reach full marks

Work out both E-cell values (-0.02 V and +0.35 V) - the standard data predict the OPPOSITE of what happens. Then explain the reversal: non-standard conditions (the values are very close) let the first reaction go, while kinetic hindrance stops the second.

Where marks are lost

Getting the two E-cell values the wrong way round (many quoted +0.02 V and -0.35 V), which makes the whole argument collapse. The reduction of Cu^{2+} has the negative value.

Examiner's insight

An excellent discriminator: most computed the E-cell values wrongly and so could not see that the standard prediction is the reverse of the observation.

17 (b)

Comment · [4 marks]

A student suggested the remaining Zn^{2+} ions can be precipitated by adding a LARGE excess of aqueous sodium hydroxide. Comment on this suggestion by describing the reactions as a large excess of sodium hydroxide is gradually added.

Model answer

✓ At first a white precipitate forms – zinc hydroxide, $Zn(H2O)4(OH)2$ (or $Zn(OH)2$). ✓ Equation: $Zn(H2O)6^{2+} + 2OH^- \longrightarrow Zn(H2O)4(OH)2 + 2H2O$ (or $Zn^{2+} + 2OH^- \longrightarrow Zn(OH)2$). ✓ But as a LARGE excess of sodium hydroxide is added, this precipitate then dissolves again. ✓ It dissolves because it forms the soluble zincate ion, $[Zn(OH)4]^{2-}$. So the student's plan fails: a large excess redissolves the zinc, rather than keeping it precipitated.

Mark scheme 4 marks

- ✓ a white precipitate forms / precipitate of $Zn(H2O)4(OH)2$ (or $Zn(OH)2$) (1)
- ✓ $Zn(H2O)6^{2+} + 2OH^- \longrightarrow Zn(H2O)4(OH)2 + 2H2O$ (or $Zn^{2+} + 2OH^- \longrightarrow Zn(OH)2$) (1)
- ✓ with excess NaOH the precipitate dissolves (1)
- ✓ due to formation of $[Zn(OH)4]^{2-}$ (1)

Accept: sufficient NaOH is needed first to neutralise the excess nitric acid; solid / crystals for the precipitate; correct formulae for the zincate even within an unbalanced equation

Ignore: state symbols even if incorrect; omission of square brackets; comments on the validity of the procedure

How to reach full marks

Describe both stages: a white zinc hydroxide precipitate forms first, then a LARGE excess of hydroxide redissolves it as the zincate ion $[Zn(OH)4]^{2-}$. The correct zincate formula is the mark most often missed.

Where marks are lost

Writing the wrong zincate formula - $[Zn(OH)6]^{4-}$ or $[Zn(OH)6]^{2-}$ instead of $[Zn(OH)4]^{2-}$. Everyone knows the precipitate redissolves; the exact complex is the trap.

Examiner's insight

Handled well - almost all knew the white precipitate redissolves - but the zincate formula was often given as $[Zn(OH)6]^{4-}$ rather than $[Zn(OH)4]^{2-}$.

17 (c)

Suggest · [2 marks]

Suggest why gilding metals are less malleable than pure copper, by considering their structure.

Model answer

✓ The zinc ions are a different size from the copper ions, so they disrupt the regular layers (the ordered structure) of copper ions in the lattice. ✓ With the layers disrupted, they are less able to slide over one another, so the metal is less malleable.

Mark scheme 2 marks

- ✓ zinc ions disrupt the layers / structure of copper ions (zinc ions are a different size / larger) (1)
- ✓ so the layers (of copper ions) are less able to slide over each other (1)

Accept: reference to atoms; the reverse argument (pure copper is more malleable because its layers slide easily)

Do not accept: M1 if particles are called 'molecules' or forces are called 'intermolecular forces'

How to reach full marks

Two linked ideas: the differently-sized zinc ions disrupt the ordered layers of copper ions, so those layers cannot slide over one another as easily - hence less malleable. Use 'ions' (or atoms), never 'molecules'.

Where marks are lost

Referring to 'intermolecular forces' or 'molecules'. A metal is a lattice of positive ions; malleability is about layers of ions sliding.

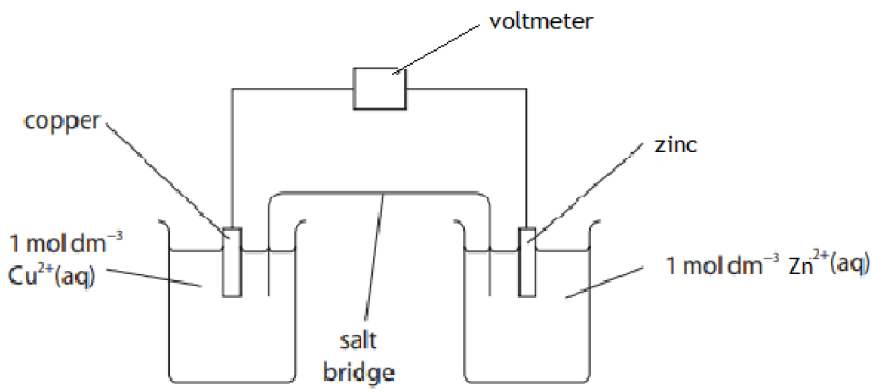
Examiner's insight

Most gained a mark for the layers not sliding; the alarming error was talk of 'intermolecular forces being disrupted' in a metal.

17 (d)(i)

Draw · [3 marks]

Draw a labelled diagram of the apparatus used to measure the emf of a cell with copper and zinc electrodes under standard conditions.



A labelled cell: a copper electrode in 1 mol dm^{-3} $Cu^{2+}(aq)$ and a zinc electrode in 1 mol dm^{-3} $Zn^{2+}(aq)$, joined by a salt bridge dipping into BOTH solutions, with a (high-resistance) voltmeter across the electrodes.

What each of the three marks needs

✓ Two correctly labelled electrodes: a copper electrode and a zinc electrode (as named in the question - not platinum). ✓ Both solutions correct with their concentrations shown: 1 mol dm^{-3} $Cu^{2+}(aq)$ and 1 mol dm^{-3} $Zn^{2+}(aq)$. ✓ A salt bridge, labelled, DIPPING INTO both solutions, with a voltmeter connected across the two electrodes.

Mark scheme 3 marks

- ✓ two correctly labelled electrodes (copper and zinc) (1)
- ✓ both solutions and their concentrations correct (1 mol dm^{-3}) (1)
- ✓ salt bridge labelled and touching both solutions, plus a voltmeter shown (1)

Accept: any soluble zinc and copper salts; names or formulae for the solutions; if the salt-bridge solution is named it must be correct

Ignore: temperature and pressure

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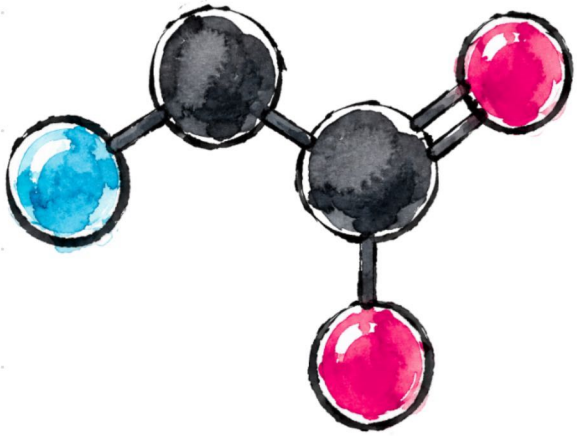

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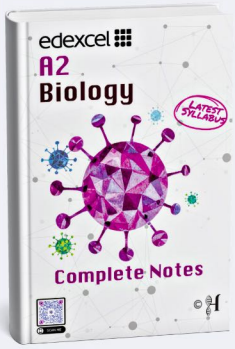

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