Electrolysis, Electrode Potentials & Cells
Mark Scheme 1

Level	International A Level
Subject	Chemistry
Exam Board	CIE
Topic	Electrochemistry
Sub-Topic	Electrolysis, Electrode Potentials & Cells
Paper Type	Theory
Booklet	Mark Scheme 1

Time Allowed: **76 minutes**

Score: **/63**

Percentage: **/100**

Grade Boundaries:

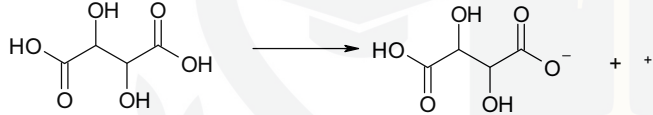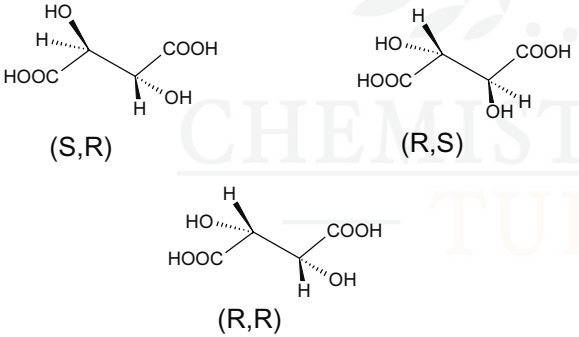| A*	A	B	C	D	E	U
>85%	777.5%	70%	62.5%	57.5%	45%	<45%

1 (a (i) m. pt. is high(er)/large(r)/greater (for iron) [1]
density is high(er)/large(r)/greater (for iron) [1]

(ii) (higher m. pt. due to)
strong attraction between cations and electrons *or*
more delocalised electrons [1]

(higher density due to) greater A_r **and** smaller radius [1]

(b) (i) components to be added: voltmeter *or* **V** [1]
salt bridge [<u>must be</u> labelled] [1]

(ii) M1:   **A** and **B**   copper (metal) or Cu **and** iron (metal) or Fe [1]
M2:   either **C** or **D**   as 1 mol dm^{-3}/1 M [1]
M3   **C** and **D** Cu^{2+} or $CuSO_4$ or $CuCl_2$ or Cu $(NO_3)_2$ etc. **and**
                        Fe^{2+} or $FeSO_4$ etc. [1]

(iii) $E^\ominus_{cell}$ = 0.34 + 0.44 = **0.78** (V) [1]

(iv) if **C** is Fe^{2+}; (as [**C**] increases), the *E* of the Fe^{2+}/Fe increases/becomes more positive/
less negative [1]

so the overall cell potential/E_{cell} would **decrease/become less positive/more
negative** [1]

or

if **C** is Cu^{2+}; (as [**C**] increases), the E of the Cu^{2+}/Cu increases/becomes more
positive/less negative [1]

so the overall cell potential/E_{cell} would **increase/become more positive/less negative**
[1]

(c) (i) (colour change is) colourless to pink/pale purple
or (end point is the first) permanent (pale) pink/pale purple colour [1]

(ii) {n(MnO_4^-) = 0.02 × 18.1/1000 = 3.62 × 10^{-4} mol}
n(Fe^{2+}) = 5 × n(MnO_4^-) = **1.81 × 10^{-3} mol** [1]

mass of Fe = 55.8 x 1.81 × 10^{-3} = 0.101 g (M2 × 55.8) ecf [1]

M_r = mass/moles = 0.500/1.81 × 10^{-3} = **276.2** ecf

[Total: 16]

2 (a A: voltmeter *or* V *or* potentiometer [1]

B: platinum *or* Pt [1]

C: $1 \, mol \, dm^{-3}$ **and** H^+ *or* HCl (*or* $0.5 \, M \, H_2SO_4$) [1]

D: lead (metal) *or* Pb [1]

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(b) a ✓ in the box next to **−0.17 V** [1]
a comment that the $[Pb^{2+}]$ has decreased **plus** a description of the outcome, e.g. as $[Pb^{2+}]$ decreases (from $1 \, mol \, dm^{-3}$), $Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$ goes over to the left hand side, *or* as $[Pb^{2+}]$ decreases, Pb^{2+} is less likely to be reduced [1]

(ii) $(K_{sp} =) [Pb^{2+}][Cl^-]^2$ [1]

(iii) if $[PbCl_2] = 3.5 \times 10^{-2}$, $[Pb^{2+}] = 3.5 \times 10^{-2}$ and $[Cl^-] = 7.0 \times 10^{-2}$
so $K_{sp} = (3.5 \times 10^{-2}) \times (7.0 \times 10^{-2})^2 = \mathbf{1.715 \, (1.7) \times 10^{-4}} \, mol^3 \, dm^{-9}$ (⩾2sf) [1]

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(c) (i) the (M^{2+}/M) $E^{\ominus}$ for the two elements are very similar *or* are −0.13 and −0.14 V [1

$E^{\ominus} (Sn^{4+}/Sn^{2+}) = 0.15 \, V$ **and** $E^{\ominus} (Pb^{4+}/Pb^{2+}) = 1.69 \, V$ [1]

so Sn^{2+} is quite easily oxidised (to Sn^{4+}) *or* is a stronger reductant *or* Pb^{2+} is not easily oxidised (to Pb^{4+}) *or* Pb^{4+} is a stronger oxidant *or* Pb^{4+} is easily reduced [1]

(ii) e.g. $PbCl_2 + Zn \longrightarrow Pb + ZnCl_2$ (*or* ionic) [1]
(other acceptable reductants: Fe, Mg, Ca but **not** Na or K)

$Sn^{2+} + Br_2 \longrightarrow Sn^{4+} + 2Br^-$ [1]
(other acceptable oxidants: VO^{2+}, $Cr_2O_7^{2-}$, Ag^+, Cl_2, Br_2, F_2, Fe^{3+}, MnO_4^-)

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(d) $Pb^{2+}(g) + 2Cl^-(g) \longrightarrow PbCl_2(s)$ [1]

(ii) $\Delta H_f = \Delta H_{at} + E(Cl - Cl) + 1^{st} \, IE + 2^{nd} \, IE + 2 \times E_A(Cl) + \text{LE}$
$-359 = 195 + \quad 242 \quad + \quad 716 + 1450 - 2 \times 349 + \quad \text{LE}$
$\text{LE} = 2 \times 349 - 359 - 195 - 242 - 716 - 1450$
$\text{LE} = \mathbf{-2264} \, (kJ \, mol^{-1})$ [3

(iii) $LE(PbCl_2) > LE(PbBr_2)$ *or* more exothermic *or* stronger lattice [1]

because Cl^- / chloride anion has smaller radius / size than Br^- / bromide [1]

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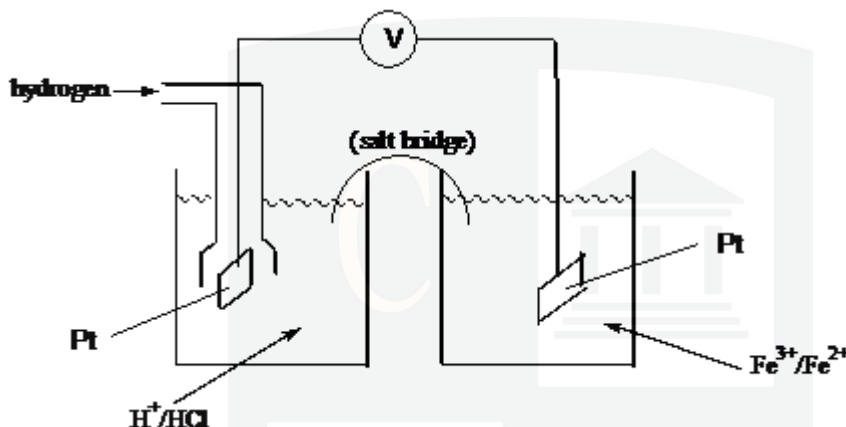**[Total: 20]**

3	(a)	(HC*l*) strong**er** acid / more dissociated / ionised in solution (HC*l* has) more ions / higher concentration of ions	1 1	**[2]**
	(b) (i)	A solution that resists changes in the pH / keeps pH *fairly* constant when **small** quantities / amounts / vols of acid / H$^+$ or base / OH$^-$ are added	1 1	
	(ii)	add (ethanoic acid) to NaOH **OR** an equation excess (ethanoic acid) **OR** mix with sodium ethanoate	1 1	**[4]**
	(c)	$CH_3CH(NH_2)COOH + H^+ \rightarrow CH_3CH(NH_3^+)COOH$ $CH_3CH(NH_2)COOH + OH^- \rightarrow CH_3CH(NH_2)COO^- + H_2O$	1 1	**[2]**
	(d) (i)		1 1	
	(ii)	any two of the above	2	**[4**

4 **(a** The potential of an **electrode** compared to that of a standard hydrogen electrode (SHE)
or
the EMF of a **cell** composed of the test electrode and the SHE [1]

all measurement concentrations of 1 mol dm^{-3} **and** 298 K / 1 atm pressure [1]
[2]

(b)

H$_2$ and good delivery system [1]
Fe^{2+}/Fe^{3+} solution labelled [1]
platinum electrodes (both) [1]
salt bridge and voltmeter [1]
H$^+$ *or* HC*l* *or* H$_2$SO$_4$ [1]
(acid is not sufficient)
[5]

(c) ($E^\ominus$ = 0.77 – 0.54 = 0.23 (V) [1]

(ii) Since $E^\ominus$ is positive/ $E^\ominus$ >0

So more products / the equilibrium will be over to the right / forward reaction is favoured
ecf from **(c)(i)** [1]

(iii) K_c = [Fe^{2+}]^2[I$_2$] / [Fe^{3+}]^2[I$^-$]^2 [1]

units are **mol^{-1} dm^3** ecf on expression [1]

(iv) ([Fe^{2+}] must always be twice [I$_2$], so) [Fe^{2+}] = 0.02 (mol dm^{-3}) [1

([I$^-$] must always be equal to [Fe^{3+}], so) [I$^-$] = 2 × 10^{-4} (mol dm^{-3}) [1

(v) K_c = {(0.02)^2 × 0.01} / {(2 x 10^{-4})^2 × (2 × 10^{-4})^2} correct expression [1]
(allow ecf from incorrect expression in **(c)(iii)**)
(allow ecf from **(c)(iv)**)
= (4 × 10^{-6}) / (1.6 × 10$^{-1.5}$) = 2.5 × 10^9 (mol^{-1} dm^3) [1
[8]

[Total: 15]