

Assessment Schedule – 2017

Chemistry: Demonstrate understanding of equilibrium principles in aqueous systems (91392)

Evidence Statement

Q	Evidence	Achievement	Merit	Excellence
ONE (a)(i)	$\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{F}^- + \text{H}_3\text{O}^+$ $\text{HBr} + \text{H}_2\text{O} \rightarrow \text{Br}^- + \text{H}_3\text{O}^+$	<ul style="list-style-type: none"> ONE correct equation. 		
(ii)	<p>To conduct a current, a substance requires mobile charged particles / ions. HF and HBr solutions have ions in solution, so both will conduct electricity. HBr is a strong acid, and therefore completely dissociates to produce a high $[\text{Br}^-]$ and $[\text{H}_3\text{O}^+]$. In contrast, HF is a weak acid and therefore only partially dissociates to produce a lower [ions], i.e. F^- and H_3O^+. So, HF is a poorer electrical conductor / weaker electrolyte.</p>	<ul style="list-style-type: none"> Mobile charged particles / ions. Relates electrical conductivity to the degree of dissociation / ion concentration. TWO species related to concentration. 	<ul style="list-style-type: none"> Links the degree of dissociation to the amount of ions for both HBr and HF. <i>(Correct follow on accepted from incorrect equations)</i> 	<ul style="list-style-type: none"> Fully compares and contrasts the electrical conductivity of HBr and HF, i.e. links the degree of dissociation to the amount of ions (not only H_3O^+), and relates this to conductivity.
(b)(i)	$\text{AgBr} \rightleftharpoons \text{Ag}^+ + \text{Br}^-$	<ul style="list-style-type: none"> Correct equation. 		
(ii)	<p>Br^- is a common ion / added to solution An increase in $[\text{Br}^-]$ will result in the reverse reaction being favoured, to restore equilibrium / minimise the change. This results in a decrease in $[\text{Ag}^+]$ (since Ag^+ reacts with some of the added Br^- to help use it up) / AgBr becomes less soluble, (until $[\text{Ag}^+][\text{Br}^-]$ is again equal to K_s.)</p>	<ul style="list-style-type: none"> Solubility of AgBr or $[\text{Ag}^+]$ decreases. 	<ul style="list-style-type: none"> Links increase in $[\text{Br}^-]$ to reverse reaction being favoured OR a decrease in solubility of AgBr OR $[\text{Ag}^+]$ decreasing. 	<ul style="list-style-type: none"> Full explanation linking an increase in $[\text{Br}^-]$ to a shift in equilibrium towards the reactants, and subsequent decrease in the solubility of AgBr OR $[\text{Ag}^+]$ decreasing.
(iii)	$\text{AgBr} \rightleftharpoons \text{Ag}^+ + \text{Br}^-$ $K_s(\text{AgBr}) = [\text{Ag}^+][\text{Br}^-]$ $5 \times 10^{-13} = [\text{Ag}^+] \times \frac{0.150 \times 40}{65}$ $[\text{Ag}^+] = 5.42 \times 10^{-12} \text{ mol L}^{-1}$	<ul style="list-style-type: none"> Correct K_s expression. 	<ul style="list-style-type: none"> Correct calculation 	<ul style="list-style-type: none"> Correct calculation with unit and significant figures.

NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response; no relevant evidence	1a	2a	4a	5a	2m	3m	2e	3e

Q	Evidence	Achievement	Merit	Excellence
TWO (a)(i)	$[\text{H}_3\text{O}^+] = \sqrt{K_a \times K_w \div [\text{NH}_3]}$ $[\text{H}_3\text{O}^+] = \sqrt{5.75 \times 10^{-10} \times 1.00 \times 10^{-14} \div 0.105}$ $[\text{H}_3\text{O}^+] = 7.40 \times 10^{-12} \text{ molL}^{-1}$ $\text{pH} = -\log(7.40 \times 10^{-12}) = 11.1$	<ul style="list-style-type: none"> • One step correct. 	<ul style="list-style-type: none"> • Correct answer, with minor error e.g. significant figures. 	<ul style="list-style-type: none"> • Correct answer, including three significant figures.
(ii)	$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$ $K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$ $10^{-9.24} = 5 \times \frac{[\text{H}_3\text{O}^+]}{1}$ $[\text{H}_3\text{O}^+] = 1.15 \times 10^{-10}$ $\text{pH} = 9.94$ $\text{p}K_a = 9.24$ <p>The solution will function as a buffer (<i>as it contains a weak base and conjugate acid in a 5:1 ratio / within a 1:10 ratio / within a pH range of 8.24 – 10.24</i>).</p> <p>Since the $[\text{NH}_3] > [\text{NH}_4^+] / \text{pH} > \text{p}K_a$, the buffer will be more effective at neutralising added strong acid.</p>	<ul style="list-style-type: none"> • Correct process for determining pH. OR Buffer use / concept 	<ul style="list-style-type: none"> • Correct pH. OR Correct evaluation of the buffer. 	<ul style="list-style-type: none"> • Correct pH, including three significant figures. AND Correct evaluation of the buffer.

<p>(b)(i)</p> <p>(ii)</p> <p>(iii)</p> <p>(c)</p>	<p>$\text{Cu(OH)}_2 \rightleftharpoons \text{Cu}^{2+} + 2\text{OH}^-$</p> <p>$K_s = [\text{Cu}^{2+}][\text{OH}^-]^2$</p> <p>Let s = solubility $[\text{Cu}^{2+}] = s$ $[\text{OH}^-]^2 = (2s)^2$ $K_s = 4s^3$ $4s^3 = 4.80 \times 10^{-20}$ So; $s = 2.29 \times 10^{-7} \text{ mol L}^{-1}$</p> <p>When copper(II) hydroxide is dissolved in an acidic solution, the H_3O^+ ions neutralise the OH^- ions / $\text{H}_3\text{O}^+ + \text{OH}^- \rightarrow 2\text{H}_2\text{O}$</p> <p>A decrease in $[\text{OH}^-]$ will result in the forward reaction being favoured, to restore equilibrium / minimise the change. This causes more solid Cu(OH)_2 to dissolve, i.e. the solubility of Cu(OH)_2 increases / so that $[\text{Cu}^{2+}][\text{OH}^-]$ will again equal K_s.</p>	<ul style="list-style-type: none"> • (i) or (ii) correct. • Correct process for calculating solubility. • Recognises OH^- ions are reacting with acid. 	<ul style="list-style-type: none"> • Correct answer for solubility, with appropriate unit. • Partial explanation for an increase in solubility. 	<ul style="list-style-type: none"> • Complete explanation for the increased solubility of Cu(OH)_2 in an acidic solution.
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NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response; no relevant evidence	1a	2a	3a	4a	3m	4m	1e	3e

(c)(i)	CH_3COO^- , Na^+ , CH_3COOH , OH^- , H_3O^+	<ul style="list-style-type: none"> • THREE species. • Both are basic • HCOOH is a stronger acid than CH_3COOH OR HCOO^- is weaker than CH_3COO^- 	<ul style="list-style-type: none"> • Links equivalence point pH for either, to the relative strength of the conjugate base OR relative concentrations of hydroxide ions. 	<ul style="list-style-type: none"> • Compares and contrasts the equivalence point for both titrations.
(ii)	<p>Both titrations produce a basic salt at the equivalence point because $[\text{OH}^-] > [\text{H}_3\text{O}^+]$ / $\text{pH} > 7$ / equations.</p> <p>At the equivalence point the ethanoic acid is converted into the ethanoate ion and the methanoic acid is converted into the methanoate ion.</p> <p>The methanoic acid titration has a lower pH than the ethanoic acid titration at the equivalence point because: methanoate ions are a weaker base / produce fewer OH^- ions than ethanoate ions OR HCOOH has a larger K_a than CH_3COOH, as it is a stronger acid. As a result, its conjugate base, HCOO^-, is weaker than CH_3COO^-, so the pH will be lower at the equivalence point.</p>			

NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response; no relevant evidence	1a	2a	3a	4a	2m	3m	2e (with minor error / omission)	2e

Cut Scores

Not Achieved	Achievement	Achievement with Merit	Achievement with Excellence
0 – 7	8 – 13	14 – 19	20 – 24