Assessment Schedule – 2018

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

Evidence Statement

Q	Evidence	Achievement	Merit	Excellence
ONE (a)	$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ $H_3O^+ = \sqrt{\dfrac{K_a \times K_w}{[CH_3COO^-]}}$ $= \sqrt{\left(\dfrac{1.74 \times 10^{-5} \times 1 \times 10^{-14}}{0.420}\right)}$ $= 6.44 \times 10^{-10}$ $pH = -\log 6.44 \times 10^{-10}$ $= 9.19$	• Correct process to determine pH.	• Correct answer, including significant figures.	
(b)(i) (ii)	$CaF_2 \rightleftharpoons Ca^{2+} + 2F^-$ $K_s = [Ca^{2+}][F^-]^2$ Let solubility be 's': $[Ca^{2+}] = s$ $[F^-] = 2s$ $K_s = 4s^3$ $s = \sqrt[3]{\left(\dfrac{3.20 \times 10^{-11}}{4}\right)}$ $= 2.00 \times 10^{-4}$ mol L^{-1}	• Correct equilibrium equation. • Method correct for determining solubility.	• Correct solubility of CaF_2 calculated, including unit, including significant figures.	

(c)(i)	In a saturated solution: $Fe(OH)_3 \rightleftharpoons Fe^{3+} + 3OH^-$ As the pH is lowered, $[H_3O^+]$ increases. The H_3O^+ will remove and neutralise OH^- / $H_3O^+ + OH^- \rightarrow 2H_2O$ A decrease in $[OH^-]$ will result in the forward reaction being favoured, to restore equilibrium / minimise the change. This causes more solid $Fe(OH)_3$ to dissolve, i.e. the solubility of $Fe(OH)_3$ increases / so that $[Fe^{3+}][OH^-]$ will again equal K_s.	• Recognises $Fe(OH)_3$ is more soluble when the pH is lowered.	• Explains that the solubility of $Fe(OH)_3$ increases due to removal of OH^- from the equilibrium.	• Fully explains, using equilibrium principles, how the solubility of $Fe(OH)_3$ increases when the pH is lowered. Must include neutralisation equation.
(ii)	The SCN^- ions can form a complex ion with Fe^{3+} ions: $Fe^{3+} + SCN^- \rightarrow [FeSCN]^{2+}$ Since the Fe^{3+} are removed from the equilibrium, more $Fe(OH)_3$ dissolves to replace some of the Fe^{3+}, i.e. equilibrium will shift towards the products / speed up the forward reaction. As a result, the solubility of $Fe(OH)_3$ increases.	• Identifies that the solubility of $Fe(OH)_3$ will increase when KSCN is added.	• Explains that the solubility of $Fe(OH)_3$ will increase due to removal of Fe^{3+} from the equilibrium / formation of a complex ion.	• Fully explains, using equilibrium principles, how the solubility of the $Fe(OH)_3$ solid increases when KSCN is added. Must include complex ion equation.

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

Q	Evidence	Achievement	Merit	Excellence
TWO (a)(i) (ii)	Volume from curve at pH 10.0 = 15 – 16.0 mL As the HCl is added, the H_3O^+ ions are neutralised by the basic component of the buffer, CH_3NH_2, according to the equation below: $CH_3NH_2 + H_3O^+ \rightarrow CH_3NH_3^+ + H_2O$ OR $CH_3NH_2 + HCl \rightarrow CH_3NH_3^+ + Cl^-$ Since the H_3O^+ are removed from the solution, the pH of the solution does not significantly change.	• Correct volume measured from curve (15 – 16 mL). • Identifies the acid / H_3O^+ is removed / neutralised by the buffer.	• Explains that the CH_3NH_2 neutralises the H_3O^+, with supporting equation.	
(b)(i) (ii)	$Cl^- > CH_3NH_3^+ > CH_3NH_2 = H_3O^+ > OH^-$ $CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$ $$K_a = 2.51 \times 10^{-11} = \frac{\left[H_3O^+\right]^2}{\left(0.168 \times \dfrac{25}{45}\right)}$$ $[H_3O^+] = 1.53 \times 10^{-6}$ mol L^{-1} pH = $-\log 1.53 \times 10^{-6}$ = 5.82	• Four correct species. • Recognises the solution at the equivalence point is acidic.	• Correct order for all. • Correct process to determine pH, but minor error, e.g. incorrect dilution of $CH_3NH_3^+$.	• Correct pH, including significant figures.
(c)	Electrical conductivity in solution requires ions. The initial solution is the weak base, methanamine. It only partially dissociates to produce a **lower [ions]**, i.e. $CH_3NH_3^+$ and OH^- ions. This means there is mainly CH_3NH_2 present, which has no charge. The solution is therefore a poor conductor of electricity compared to the solution at the equivalence point. $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$ The solution at the equivalence point is made up of the products from the titration / $CH_3NH_3^+$ and Cl^-. As there are more ions in solution / higher [ions] the solution at the equivalence point is a better conductor.	• Recognises electrical conductivity in solution requires **ions**.	• Links the identified [ions] to the electrical conductivity of ONE solution.	• Full explanation of the electrical conductivity of BOTH solutions, i.e. links [ions] to electrical conductivity, including supporting equation.

NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response; no relevant evidence.	1a	2a	3a	4a	2m	3m	1e + 1m	2e, with minor error or omission.

Q	Evidence	Achievement	Merit	Excellence
THREE (a)(i)	$K_s = [Ag^+][Cl^-]$	• Correct K_s expression.		
(ii)	The Ag^+ is a common ion. By increasing $[Ag^+]$, the equilibrium will shift towards the reactants to use up some of the extra Ag^+, so more solid AgCl will form, i.e. the solubility of AgCl will decrease.	• Identifies Ag^+ is a common ion / $[Ag^+]$ increases.	• Explains the effect of increasing $[Ag^+]$ using equilibrium principles.	
(iii)	Final $[Ag^+] = \dfrac{70}{120} \times 0.0220 = 0.0128$ mol L^{-1} Final $[Cl^-] = \dfrac{50}{120} \times 0.0550 = 0.0229$ mol L^{-1} $Q = [Ag^+][Cl^-] = 0.0128 \times 0.0229 = 2.94 \times 10^{-4}$ Since $Q > K_s$, AgCl will form a precipitate.	• Correct substitution into Q expression. OR Correct $[Ag^+]$ or $[Cl^-]$.	• Correct process to determine Q and compare with K_s.	• Correct calculation and comparison with K_s to determine whether AgCl will form a precipitate.
(b)(i)	The solution will function as a buffer over a pH range 2.74 – 4.74 ($pK_a + / - 1$).	• Identifies pH range for buffer.		
(ii)	$n(HCOONa) = \dfrac{m}{M} = \dfrac{5.11}{68} = 0.0751$ mol $c(HCOO^-) = \dfrac{n}{V} = \dfrac{0.0751}{0.125} = 0.601$ mol L^{-1} $K_a = 1.82 \times 10^{-4} = \dfrac{\left[HCOO^-\right]\left[H_3O^+\right]}{\left[HCOOH\right]}$ $= \dfrac{0.601 \times \left[H_3O^+\right]}{0.105}$ $[H_3O^+] = 3.18 \times 10^{-5}$ mol L^{-1} $pH = -\log [H_3O^+] = 4.50$	• Determines $n(HCOO^-)$. OR Correct substitution into K_a expression.	• Correct pH.	• Correct calculation of the buffer's pH. AND A full evaluation of the ability of the solution to function as a buffer.
(iii)	Since the pH of the solution falls within the buffer zone (2.74 – 4.74), it will function as a buffer. However, as the pH > pK_a, / this means $[HCOO^-] > [HCOOH]$, so the buffer will be more effective against added strong acid.	• Buffer is more effective against acid.	• Evaluates ability of solution to function as a buffer (could have an incorrect pH).	

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No response; no relevant evidence.	1a	2a	3a	4a	2m	3m	2e, with minor error or omission.	2e

Cut Scores

Not Achieved	Achievement	Achievement with Merit	Achievement with Excellence
0 – 8	9 – 13	14 – 19	20 – 24