Assessment Schedule – 2022 FINAL

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

Evidence Statement

Q	Evidence	Achievement	Merit	Excellence
ONE (a)(i)	$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{2-}$	• Correct equilibrium equation and K_s expression.		
(ii)	$K_s = [Ag^+]^2[CrO_4^{2-}]$			
(iii)	$K_s = (2 \times 6.50 \times 10^{-5})^2 \times (6.50 \times 10^{-5})$ $= 1.10 \times 10^{-12}$	• Method correct for determining K_s.	• Correct K_s (accept 2 – 4 significant figures).	
(b)(i)	When $AgNO_3$ is added, $[Ag^+]$ increases. Since Ag^+ is a common ion, this causes the reverse reaction of the equilibrium to speed up to use up some of the extra Ag^+. As a result, more solid Ag_2CrO_4 is formed, i.e. the solubility of Ag_2CrO_4 decreases.	• Recognises $[Ag^+]$ increases / Ag^+ is a common ion.	• Explains that increase in $[Ag^+]$ favours the reverse reaction.	
(ii)	When NaCN is added, the CN^- ions form a complex ion with the Ag+ ions: $Ag^+ + 2CN^- \rightarrow [Ag(CN)_2]^-$ Since the Ag^+ ions are removed from the equilibrium, the forward reaction speeds up to replace some of the Ag^+. This causes more solid Ag_2CrO_4 to dissolve, i.e. the solubility of Ag_2CrO_4 increases.	• Recognises Ag^+ ions form a complex ion with CN^- ion OR • Ag^+ reacts with CN^- so $[Ag^+]$ decreases.	• Explains that the formation of a complex ion removes Ag^+ from equilibrium / causes decrease in $[Ag^+]$.	• Fully explains, using equilibrium principles, how the solubility of Ag_2CrO_4 is affected by addition of a common ion and formation of a complex ion. Complex ion equation required.
(c)	$IP = \left(\dfrac{50}{80} \times 0.00154\right) \times \left(2 \times \dfrac{30}{80} \times 0.0686\right)$ $= 4.95 \times 10^{-5}$ Since IP > K_s, a precipitate of AgCl will form.	• EITHER correct $[Ag^+]$ OR $[Cl^-]$ • Correct comparison of IP with K_s.	• Correct process to calculate IP and compare with K_s, but does not multiply $[Cl^-]$ by 2.	• Correct calculation (accept 2 – 4 significant figures) and comparison with K_s to show that AgCl will form a precipitate.

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

Q	Evidence	Achievement	Merit	Excellence
TWO (a)(i)	$Na^+ > OCl^- > HOCl = OH^- > H_3O^+$	• THREE correct species identified in the correct order.		
(ii)	$HOCl + H_2O \rightleftharpoons OCl^- + H_3O^+$ $$K_a = \frac{\left[OCl^-\right]\left[H_3O^+\right]}{\left[HOCl\right]}$$ $$2.95 \times 10^{-8} = \frac{\left(10^{-4.80}\right)^2}{\left[HOCl\right]}$$ $$\left[HOCl\right] = 0.00851 \, mol \, L^{-1}$$	• Correct process for calculating [HOCl].	• Correctly calculates [HOCl], including correct unit (accept 2 – 4 significant figures).	
(iii)	NaOCl is a basic salt and completely dissociates to produce a high [ions], i.e. Na^+ and OCl^- ions, in solution. $NaOCl \rightarrow Na^+ + OCl^-$ Since the [ions] is high, NaOCl solution will be a good electrical conductor. HOCl is a weak acid and only partially dissociates to produce a relatively low [ions], i.e. OCl^- and H_3O^+ ions, in solution. $HOCl + H_2O \rightleftharpoons OCl^- + H_3O^+$ Since the [ions] is low, HOCl solution will be a poor electrical conductor.	• Recognises a solution requires mobile ions for electrical conductivity.	• Explains why NaOCl is a good electrical conductor. OR Explains why HOCl is a poor electrical conductor.	• Fully compares the electrical conductivity of BOTH solutions, including relevant equations (equation for HOCl must include water).
(b)	When HCl is added, the H_3O^+ ions are neutralised by the basic component of the buffer, OCl^-. As a result, the pH of the solution does not significantly change, since the H_3O^+ ions are removed from the solution / HOCl produced is a much weaker acid than HCl.	• Identifies the acid / H_3O^+ ions are removed / neutralised by the buffer.	Explains that the OCl^- neutralises/removes the H_3O^+ and therefore the pH does not significantly change.	

(c)	Both HF and CH_3COOH are weak acids. Since the solutions have the same concentration, but HF has a lower pH, this means HF dissociates to a greater extent than CH_3COOH/HF must have a greater K_a than CH_3COOH. This means HF has a higher $[H_3O^+]$ and therefore a lower pH. CH_3COONa is a basic salt. The CH_3COO^- is a weak base and therefore partially dissociates to produce a high $[OH^-]$. This means CH_3COONa has a lower $[H_3O^+]$ than both HF and CH_3COOH, so it has the highest pH. $HF + H_2O \rightleftharpoons F^- + H_3O^+$ $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$	• Recognises HF has the highest $[H_3O^+]$ / CH_3COO^- has the lowest $[H_3O^+]$.	• Explains the order for TWO of the solutions (ok to refer to only OH^- for CH_3COO^-).	• Fully justifies the order in terms of $[H_3O^+]$, including at least two relevant equations.

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

Q	Evidence	Achievement	Merit	Excellence
THREE (a)	For this solution: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ $[NH_3] = 0.124$ mol L^{-1} For this solution: $[NH_4^+] = [OH^-] = \dfrac{K_w}{\left[H_3O^+\right]}$ Using K_a to calculate pH: $K_a = \dfrac{\left[NH_3\right]\left[H_3O^+\right]}{\left[NH_4^+\right]}$ $5.75 \times 10^{-10} = 0.124 \times \dfrac{\left[H_3O^+\right]^2}{10^{-14}}$ $\left[H_3O^+\right] = 6.81 \times 10^{-12}$ mol L^{-1} $pH = 11.2$	• Correct process for calculating pH of ammonia solution.	• Correctly calculates pH of ammonia solution (accept 2 – 4 significant figures).	
(b)(i)	$K_a = \dfrac{\left[NH_3\right]\left[H_3O^+\right]}{\left[NH_4^+\right]}$ $5.75 \times 10^{-10} = 5 \times \dfrac{\left[H_3O^+\right]}{2}$ $\left[H_3O^+\right] = 2.30 \times 10^{-10}$ $pH = 9.64$ Since pH > pK_a, $[NH_3] > [NH_4^+]$. This means the buffer will be more effective at neutralising small volumes of strong acid. OR Evaluates effectiveness using 5 : 2 ratio.	• Correct process for calculating pH of buffer solution. • Recognises the buffer is more effective at neutralising a strong acid.	• Correctly calculates pH of buffer solution. OR Since $[NH_3] > [NH_4^+]$ / [base] > [acid], the buffer is more effective at neutralising a strong acid / evaluates effectiveness using 5:2 ratio.	• Correct pH of buffer solution. AND Explains why the buffer is more effective at neutralising strong acid. Must compare relative $[NH_3]$ and $[NH_4^+]$ for E8. If only compares [base] and [acid], then minor error.
(ii)	The pH is 9.24 halfway to the equivalence point. Halfway to the equivalence point, $[NH_3] = [NH_4^+]$ since half of the NH_3 has reacted with HNO_3 to form NH_4^+. At this point, the pH of the solution will be equal to the pK_a since $1NH_3 : 1NH_4^+$.	• Identifies the correct pH halfway to the equivalence point/pH equals pK_a OR Recognises $[NH_3] = [NH_4^+]$.	• Explains that the pH equals the pK_a since half of the NH_3 has reacted (with HNO_3) to form NH_4^+, so $[NH_3] = [NH_4^+]$.	

(c)(i)	CH₃COOH is a stronger acid than NH₄⁺ since it has a larger K_a. As a result, the conjugate base, CH₃COO⁻, will be weaker base than NH₃ and will therefore have a lower [OH⁻] / higher [H₃O⁺]. This means the initial pH will be lower.	• Circles "lower" for both parts (i) and (ii) OR • Circles "lower" for one and correctly identifies differences in strength of either relevant bases for (i) or acids for (ii).	• Explains either (i) or (ii) by comparing relative size of K_a / pK_a.	• Explains both (i) and (ii) in terms of magnitude of K_a / pK_a, and either relative [H₃O⁺] or [OH⁻].
(ii)	At the equivalence point, the weak acids CH₃COOH and NH₄⁺ will be present. Since CH₃COOH is a stronger acid than NH₄⁺, it will have a higher [H₃O⁺] and therefore a lower pH at the equivalence point.			

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

Cut Scores

Not Achieved	Achievement	Achievement with Merit	Achievement with Excellence
0 – 7	8 – 13	14 – 18	19 – 24