Assessment Schedule – 2025

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

Evidence Statement

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
ONE (a)(i)	$Fe(OH)_2(s) \rightleftharpoons Fe^{2+}(aq) + 2OH^-(aq)$	• Correct equation and K_s expression.		
(ii)	$K_s = [Fe^{2+}][OH^-]^2$			
(iii)	$K_s = (1.01 \times 10^{-5}) \times (2 \times 1.01 \times 10^{-5})^2$ $K_s = 4.12 \times 10^{-15}$	• Method correct for determining K_s.	• Correct K_s, including 2–4 significant figures.	
(iv)	When HCl is added, the $[H_3O^+]$ increases. The H_3O^+ react with the OH^- (in an acid-base reaction). The equation for this reaction is: $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(\ell)$ This decreases $[OH^-]$ in the system, so the rate of the forward reaction increases to replace the OH^- ions. This causes more $Fe(OH)_2(s)$ to dissolve and hence $Fe(OH)_2$ becomes more soluble.	• Recognises $[H_3O^+]$ increases.	• Explains that a decrease in $[OH^-]$ due to reaction with H_3O^+ will favour formation of more OH^- in solution / favour forward reaction.	• Justifies what will happen to the solubility of $Fe(OH)_2$ when HCl is added, including relevant balanced equation. AND Explains that the formation of the complex ion causes the rate of the forward / dissolving reaction to increase to replace the Ni^{2+}, including the balanced equation to show complex ion formation.
(b)(i)	When NaCN is added, the Ni^{2+} ions form a complex with the added CN^- ions: $Ni^{2+}(aq) + 4CN^-(aq) \rightarrow [Ni(CN)_4]^{2-}(aq)$ This decreases $[Ni^{2+}]$ in the system, so the rate of the $Ni(OH)_2$ dissolving reaction increases to replace the Ni^{2+} ions. This causes the precipitate to dissolve/solubility increases. Because an excess of cyanide ions is added and the complex ion is water soluble, eventually all of the $Ni(OH)_2$ dissolves.	• Recognises the $Ni(OH)_2$ precipitate dissolves due to formation of a complex ion.	• Explains that the formation of the complex ion causes a decrease in $[Ni^{2+}]$.	

(ii)	$[Ni^{2+}]$ in mixture $= \frac{0.130 \times 55.0}{90.0}$ $= 0.0794 \text{ mol L}^{-1}$ $[OH^-]$ in mixture $= \frac{1 \times 10^{-14}}{10^{-11.7}} \times \frac{35.0}{90.0}$ $= 1.95 \times 10^{-3} \text{ mol L}^{-1}$ In the mixture: IP $= [Ni^{2+}][OH^-]^2$ $= 0.0794 \times (1.95 \times 10^{-3})^2$ $= 3.02 \times 10^{-7}$ Since IP > K_s, a precipitate forms.	• Calculates $[Ni^{2+}]$ or $[OH^-]$. • Correctly compares IP and K_s.	• Correct process to calculate IP and compare with K_s, but one of the solutions is not appropriately diluted. e.g. $[OH^-]$ given as 5.01×10^{-3} (IP is 1.99×10^{-6}). Cannot multiply $[OH^-]$ by 2.	• Correct calculation and comparison, including 2–4 significant figures.

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.	2e

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
TWO (a)	Cl^-, CH_3NH_2, OH^-	• All correct.		
(b)(i)	A buffer solution resists a change in pH / maintains a fairly constant pH when a small volume of either strong acid or strong base is added. When NaOH is added, the $CH_3NH_3^+$ ions react with (neutralise) the OH^- to form the weaker base, CH_3NH_2. As a result, the pH does not change significantly.	• Correct description of the function of a buffer solution.	• Explains how the buffer resists a change in pH when NaOH is added and that the $CH_3NH_3^+$ ions / acidic component of buffer neutralise / react with / remove the added OH^- ions.	
(ii)	$CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$ $K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$ $2.51 \times 10^{-11} = \frac{0.840 \times 10^{-12.1}}{[CH_3NH_3^+]}$ $[CH_3NH_3^+] = 0.0266 \text{ mol L}^{-1}$ $n(CH_3NH_3^+) = cv = 0.0266 \times 0.600 = 0.0159 \text{ mol}$ $m(CH_3NH_3Cl) = 0.0159 \times 67.5 = 1.08 \text{ g}$	• Correct process to calculate $[CH_3NH_3^+]$.	• Correct process to calculate the mass of CH_3NH_3Cl, but minor error.	• Correctly calculates the mass of CH_3NH_3Cl, including unit and 2–4 significant figures.
(iii)	For a solution to function as a buffer solution, the pH of the solution needs to be equal to pK_a+/–1. For this buffer solution, the pH would therefore need to fall within the range 9.6 – 11.6. Since the pH is 12.1, the solution will not function as a buffer. The pH of the solution needs to be decreased. This could be achieved by adding a strong acid, like HCl, until the pH falls within the range. Alternatively, solid CH_3NH_3Cl could be added.	• Identifies why the solution cannot function as a buffer. OR Recognises the pH of the solution needs to decrease.	• Explains why the solution cannot function as a buffer with reference to pK_a and the pH range. OR Explains how the pH could be decreased to the appropriate pH range by addition of an acid / H_3O^+ ions.	AND Explains why the solution cannot function as a buffer and outlines how the solution needs to be modified for use as a buffer solution by adding a strong acid / HCl / H_2SO_4 / HNO_3 or (solid) CH_3NH_3Cl to lower the pH.

(c)	HBr is a strong acid and completely dissociates to produce a relatively high $[H_3O^+]$. As a result, HBr has a low pH. $HBr + H_2O \rightarrow Br^- + H_3O^+$ NH_4Cl is an acidic salt. The NH_4^+ is a weak acid and partially dissociates to produce a lower $[H_3O^+]$ than HBr. As a result, NH_4^+ has a higher pH than HBr. $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ CH_3NH_2 is a weak base and partially dissociates to produce a higher $[OH^-]$ than the other two solutions, and therefore the lowest $[H_3O^+]$ of the three solutions. As a result, CH_3NH_2 has the highest pH. $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$	• Recognises HBr has the highest $[H_3O^+]$ / CH_3NH_2 has the lowest $[H_3O^+]$. • Recognised HBr fully dissociates, but NH_4^+ or CH_3NH_2 only partially dissociates.	• Explains the order for TWO of the solutions (OK to refer to only OH^- for CH_3NH_2).	• Justifies the order in terms of degree of dissociation and $[H_3O^+]$, including at least THREE 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.	2e

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
THREE (a)	A suitable indicator would need to change colour at the equivalence point/over the vertical section of the curve. From the curve, the equivalence point occurs around pH 5.8/vertical section from around 4-8. Since methyl red has a pK_a of 5.1, it will change colour over a pH range of 4.1 – 6.1, so will change colour at the equivalence point/over the vertical section. Thymol blue will change colour over a pH range of 0.7 – 2.7, so it will be unsuitable since it will change colour **after** the equivalence point.	• Recognises methyl red will change colour at the equivalence point since its p*K*a is close to the equivalence point / falls in vertical section. OR Thymol blue will change colour after the equivalence point since its p*K*a falls **after** the equivalence point / vertical section.	• Explains the suitability of each indicator in terms of the pH range over which each indicator changes colour relative to the estimated equivalence point pH.	
(b)(i)	$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ $K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$ $5.75 \times 10^{-10} = \frac{[H_3O^+]^2}{\left(0.0139 \times \frac{25}{45}\right)}$ $[H_3O^+] = 2.11 \times 10^{-6}$ mol L^{-1} $pH = -\log(2.11 \times 10^{-6})$ $= 5.68$	• Correct process to calculate pH of an acidic solution.	• Calculates pH, but does not appropriately dilute the solution (pH will be 5.55 if does not dilute).	• Correctly calculates pH at equivalence point, including 2–4 significant figures.
(ii)	The initial ammonia solution is weakly basic, so the NH_3 only partially dissociates to produce a relatively low $[NH_4^+]$ and $[OH^-]$: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ Since [ions] is low, the solution is a poor conductor. After 20.0 mL of HCl has been added, the solution in the conical flask consists of an acidic salt that has completely dissociated to produce high $[NH_4^+]$ ions and high $[Cl^-]$ ions. Although a small proportion of the NH_4^+ ions will dissociate, the solution still has a high [ions] and is therefore a good electrical conductor. $NH_4Cl \rightarrow NH_4^+ + Cl^-$	• Recognises a solution requires ions to conduct electricity. • Identifies NH_3 as a poor conductor and NH_4Cl as a good conductor. OR Recognises electrical conductivity increases.	• Links conductivity of ONE solution to the degree of dissociation and [ions] in solution.	AND Links conductivity of BOTH solutions to the degree of dissociation and [ions] in solution, including at least one supporting equation.

(c)(i)	$n(\text{HCl}) = cv = 0.0174 \times 0.005 = 8.7 \times 10^{-5}$ mol $c(\text{HCl}) = \dfrac{8.7 \times 10^{-5}\ \text{mol}}{0.05\ \text{L}} = 1.74 \times 10^{-3}$ mol L^{-1} $\text{pH} = -\log 1.74 \times 10^{-3}$ $= 2.76$	• Calculates the correct moles of HCl. OR [HCl]	• Correct process for calculating pH, but minor error. OR	• Correct pH of diluted HCl and explanation as to why the pH is higher than the original HCl solution.
(ii)	The pH is higher because the volume of solution in which the HCl is contained is much larger / the HCl has been diluted. As a result, there is a decrease in [H_3O^+], so the pH is higher.	• Recognises the HCl has been diluted / there are less H_3O^+ per unit volume.	Links increase in pH of HCl to decrease in [H_3O^+] due to dilution.	

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.	2e

Cut Scores

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