Assessment Schedule – 2020

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

Evidence Statement

Q	Evidence	Achievement	Merit	Excellence
ONE (a)(i)	8.24 – 10.24	• Correct pH range.		
(ii)	The H_3O^+ ions from the HNO_3 solution are neutralised /removed from the solution by the basic component of the buffer, NH_3, according to the equation below: $NH_3 + H_3O^+ \rightarrow NH_4^+ + H_2O$ $(HNO_3 + NH_3 \rightarrow NH_4^+ + NO_3^-)$ The pH of the solution does not significantly change since the H_3O^+ are removed from the solution / NH_4^+ produced is a much weaker acid than HNO_3.	• Identifies the acid/ H_3O^+ ions are removed / neutralised by the buffer / correct equation.	• Explains that the NH_3 neutralises the H_3O^+, with supporting equation.	
(b)(i)	$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ $K_a = \dfrac{\left[NH_3\right]\left[H_3O^+\right]}{\left[NH_4^+\right]}$ $5.75 \times 10^{-10} = \dfrac{0.0500 \times 10^{-8.75}}{\left[NH_4^+\right]}$ $\left[NH_4^+\right] = 0.155 \text{ mol L}^{-1}$ $n(NH_4Cl) = cV = 0.155 \times 0.200 = 0.0309 \text{ mol}$ $m(NH_4Cl) = n \times M = 0.0309 \times 53.5 = 1.65 \text{ g}$	• Correctly substitutes into K_a expression / other suitable equation.	• Correct process to determine mass of NH_4Cl required, with minor error.	• Correctly calculates mass of NH_4Cl required. AND Explains why the buffer is more effective at neutralising strong base.
(ii)	Since the pH of the buffer is less than the pK_a (8.75 < 9.24), $[NH_4^+] > [NH_3]$. As a result, the buffer will be more effective at neutralising small volumes of added strong base.	• Recognises the buffer is more effective at neutralising a strong base.	• Since $[NH_4^+] > [NH_3]$ / [acid] > [base], the buffer is more effective at neutralising a strong base.	

(c)(i)	Order of decreasing pH: CH_3NH_2, NH_4Cl, HBr. CH_3NH_2 is a weak base, and therefore partially dissociates to produce OH^-. $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$ NH_4Cl is an acidic salt. The NH_4^+ is a weak acid, and therefore partially dissociates to produce H_3O^+. $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ HBr is a strong acid, and therefore completely dissociates to produce a higher $[H_3O^+]$ than NH_4^+. $HBr + H_2O \rightarrow Br^- + H_3O^+$ Since CH_3NH_2 is the only base, it has the lowest $[H_3O^+]$ and therefore the highest pH. Since HBr is a strong acid, it has the highest $[H_3O^+]$ and therefore the lowest pH.	• Recognises $[H_3O^+]$ affects pH.	• Correct order, and some explanation of degree of dissociation and $[H_3O^+]$ linked to pH.	• Fully justifies order, including degree of dissociation and relative $[H_3O^+]$ linked to pH, with support from at least TWO correct equations, AND correct $[NH_4^+]$.
(ii)	$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ $$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$ $$5.75 \times 10^{-10} = \frac{(10^{-4.70})^2}{[NH_4^+]}$$ $[NH_4^+] = 0.692 \text{ mol L}^{-1}$	• Correct process for calculating $[NH_4^+]$.	• Correct $[NH_4^+]$, including unit.	

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

Q	Evidence	Achievement	Merit	Excellence
TWO (a)(i) (ii) (iii)	$PbBr_2 \rightleftharpoons Pb^{2+} + 2Br^-$ $K_s = [Pb^{2+}][Br^-]^2$ Let solubility be s: $K_s = 4s^3$ $s = \sqrt[3]{\left(\dfrac{2.10 \times 10^{-6}}{4}\right)}$ $s = 8.07 \times 10^{-3}$ mol L^{-1}	• Correct equilibrium equation and K_s expression. • Method correct for determining solubility of PbBr$_2$.	• Correct solubility of PbBr$_2$, including unit.	
(b)	$\left[Pb^{2+}\right] = \dfrac{125}{300} \times 0.0365 = 1.52 \times 10^{-2}$ mol L^{-1} $\left[Br^-\right] = \dfrac{175}{300} \times 2 \times 0.00262 = 3.06 \times 10^{-3}$ mol L^{-1} $IP = \left[Pb^{2+}\right]\left[Br^-\right]^2 = \left(1.52 \times 10^{-2}\right) \times \left(3.06 \times 10^{-3}\right)^2$ $= 1.42 \times 10^{-7}$ Since IP < K_s, no PbBr$_2$ precipitate will form.	• EITHER Correct [Pb^{2+}] OR [Br$^-$]. • Correct comparison of Q_s (IP) with K_s	• Correct process to determine Q_s (IP) and compare with K_s.	• Correct calculation and comparison with K_s to determine whether PbBr$_2$ will form a precipitate.

(c)(i)	$Ni(OH)_2 \rightleftharpoons Ni^{2+} + 2OH^-$ When ammonia solution is added, the Ni^{2+} ions are removed from the saturated solution to form a complex ion, as shown below: $Ni^{2+} + 6NH_3 \rightarrow [Ni(NH_3)_6]^{2+}$ To restore the equilibrium, more solid $Ni(OH)_2$ dissolves to replace some of the Ni^{2+} / increase $[Ni^{2+}]$. Therefore the solubility of $Ni(OH)_2$ increases. When the pH is decreased below pH 4, $[H_3O^+]$ increases. The H_3O^+ remove the OH^- from the equilibrium according to the neutralisation equation below: $H_3O^+ + OH^- \rightarrow 2H_2O$ To restore the equilibrium, more solid $Ni(OH)_2$ dissolves to replace some of the OH^- / increase $[OH^-]$. Therefore the solubility of $Ni(OH)_2$ increases.	• Recognises Ni^{2+} forms a complex ion with NH_3. OR Recognises $[H_3O^+]$ increases below pH 4.	• Explains that the formation of a complex ion removes Ni^{2+} from equilibrium / causes decrease in $[Ni^{2+}]$. • Explains that H_3O^+ removes OH^- from equilibrium / causes decrease in $[OH^-]$.	• Fully explains, using equilibrium principles, how the solubility of $Ni(OH)_2$ increases due to both the addition of NH_3 and decreasing pH below 4.
(ii)	$Ni(OH)_2 \rightleftharpoons Ni^{2+} + 2OH^-$ $K_s = \left[Ni^{2+} \right]\left[OH^- \right]^2$ $\left[OH^- \right] = \dfrac{1 \times 10^{-14}}{10^{-8.25}} = 1.78 \times 10^{-6} \text{ mol L}^{-1}$ $6.00 \times 10^{-16} = \left[Ni^{2+} \right] \times \left(1.78 \times 10^{-6} \right)^2$ $\left[Ni^{2+} \right] = 1.90 \times 10^{-4} \text{ mol L}^{-1}$	• Correctly calculates $[OH]^-$.	• Correctly calculates $[Ni^{2+}]$, i.e. solubility of $Ni(OH)_2$, including unit.	

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

Q	Evidence	Achievement	Merit	Excellence
THREE (a)(i)	Na^+, CH_3COO^-, CH_3COOH, OH^-, H_3O^+	• THREE correct species identified.		
(ii)	For this solution: $$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$$ $[CH_3COO^-] = 0.224 \text{ mol L}^{-1}$ For this solution: $$\left[CH_3COOH\right] = \left[OH^-\right] = \frac{K_w}{\left[H_3O^+\right]}$$ Using K_a to calculate pH: $$K_a = \frac{\left[H_3O^+\right]\left[CH_3COO^-\right]}{\left[CH_3COOH\right]}$$ $$1.74 \times 10^{-5} = \frac{0.224 \times \left[H_3O^+\right]^2}{10^{-14}}$$ $$\left[H_3O^+\right] = 8.81 \times 10^{-10} \text{ mol L}^{-1}$$ $$pH = 9.05$$	• Correct process for determining pH of sodium ethanoate solution.	• Correctly calculates pH of sodium ethanoate solution.	
(b)	The electrical conductivity of a solution depends upon its concentration of mobile ions. HCl is a strong acid and therefore completely dissociates (evidence from equation) to produce a high concentration of Cl^- ions and H_3O^+ ions in solution. $$HCl + H_2O \rightarrow Cl^- + H_3O^+$$ Sodium ethanoate is a basic salt. It completely dissociates (evidence from equation) to produce a high concentration of Na^+ ions and CH_3COO^- ions in solution. $$CH_3COONa \rightarrow Na^+ + CH_3COO^-$$ Since both solutions produce a high concentration of ions, they are both good electrical conductors.	• Recognises a solution requires mobile ions for electrical conductivity.	• Explains why HCl is a good electrical conductor. OR Explains why sodium ethanoate is a good electrical conductor.	Fully justifies why BOTH solutions are good electrical conductors, including equations.

(c)(i)	$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ $K_a = \dfrac{\left[CH_3COO^-\right]\left[H_3O^+\right]}{\left[CH_3COOH\right]}$ $1.74 \times 10^{-5} = \dfrac{\left[H_3O^+\right]^2}{\left(0.224 \times \frac{25}{45}\right)}$ $\left[H_3O^+\right] = 1.47 \times 10^{-3}$ mol L^{-1} pH = 2.83	• Correct process for calculating pH at equivalence point.	• Calculates correct pH at equivalence point.	• Calculates correct pH at equivalence point AND Explains why the pH is higher at the equivalence point for (ii).
(ii)	Circled "Higher pH". At the equivalence point, the weak acid is present, i.e. either $CH_3NH_3^+$ or CH_3COOH. Since $CH_3NH_3^+$ has a smaller K_a (larger pK_a), it is a weaker acid than CH_3COOH and will therefore dissociate to a lesser extent to produce a lower $[H_3O^+]$. As a result, the pH at the equivalence point will be higher for $CH_3NH_3^+$.	• Circles "Higher pH".	• Links the pH at the equivalence point to the pH of the weak acid and the magnitude of its K_a / pK_a.	

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