

Evidence Statement

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
ONE (a)(i)	$\text{Ag}_2\text{SO}_4 \rightleftharpoons 2\text{Ag}^+ + \text{SO}_4^{2-}$	<ul style="list-style-type: none"> Correct equilibrium equation and K_s expression. 		
(ii)	$K_s = [\text{Ag}^+]^2[\text{SO}_4^{2-}]$			
(iii)	Let solubility be s : $[\text{Ag}^+] = 2s$ $[\text{SO}_4^{2-}] = s$ $K_s = 1.20 \times 10^{-5} = 4s^3$ $s = 0.0144 \text{ mol L}^{-1}$ $[\text{Ag}^+] = 0.0288 \text{ mol L}^{-1}$ $[\text{SO}_4^{2-}] = 0.0144 \text{ mol L}^{-1}$	<ul style="list-style-type: none"> Correct process for determining solubility. 	<ul style="list-style-type: none"> Correct solubility, and $[\text{Ag}^+]$ and $[\text{SO}_4^{2-}]$, including unit and significant figures (2-4). 	
(b)(i)	Add NH_3 . The NH_3 reacts with the Ag^+ from the equilibrium to form a complex ion: $\text{Ag}^+ + 2\text{NH}_3 \rightarrow [\text{Ag}(\text{NH}_3)_2]^+$ This causes a decrease in $[\text{Ag}^+]$. As a result, the rate of the forward reaction speeds up to replace some of the Ag^+ . This causes more solid Ag_2SO_4 to dissolve, so the solubility increases .	<ul style="list-style-type: none"> Recognises Ag^+ can form a complex ion with NH_3. 	<ul style="list-style-type: none"> Explains that the formation of a complex ion removes Ag^+ from equilibrium / causes decrease in $[\text{Ag}^+]$. 	<ul style="list-style-type: none"> Uses equilibrium principles to justify the choice of solution to change solubility, including equation for complex ion formation.
(ii)	Add Na_2SO_4 . The SO_4^{2-} ion is a common ion. The increase in $[\text{SO}_4^{2-}]$ causes the rate of the reverse reaction to speed up to use up some of the extra SO_4^{2-} . This produces more solid Ag_2SO_4 , so the solubility decreases.	<ul style="list-style-type: none"> Recognises SO_4^{2-} is a common ion / $[\text{SO}_4^{2-}]$ increases. 	<ul style="list-style-type: none"> Explains that increase in $[\text{SO}_4^{2-}]$ favours the reverse reaction. 	

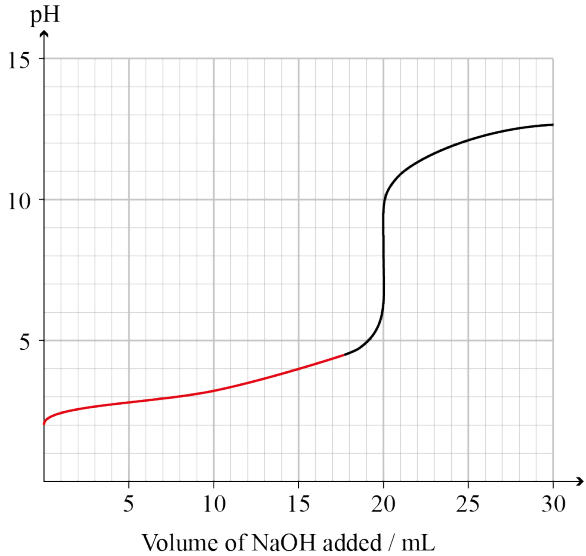
(c)	$[\text{Ag}^+] = \frac{20}{50} \times 0.0188 = 7.52 \times 10^{-3} \text{ mol L}^{-1}$ $[\text{SO}_4^{2-}] = \frac{30}{50} \times 3 \times 0.0146 = 0.02628 \text{ mol L}^{-1}$ $\text{IP} = [\text{Ag}^+]^2[\text{SO}_4^{2-}]$ $= (7.52 \times 10^{-3})^2 \times (0.02628)$ $= 1.49 \times 10^{-6}$ <p>Since $\text{IP} < K_s$, a precipitate of Ag_2SO_4 will not form.</p>	<ul style="list-style-type: none"> • EITHER correct $[\text{Ag}^+]$ OR $[\text{SO}_4^{2-}]$. • Correct comparison of IP with K_s. 	<ul style="list-style-type: none"> • Correct process to calculate IP and compare with K_s but does not multiply $[\text{SO}_4^{2-}]$ by 3. 	<ul style="list-style-type: none"> • Correct calculation (accept 2 – 4 significant figures), and comparison with K_s, to show that Ag_2SO_4 will not form a precipitate.
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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)(i)	$\text{Na}^+ > \text{CH}_3\text{COO}^- > \text{CH}_3\text{COOH} = \text{OH}^- > \text{H}_3\text{O}^+$	<ul style="list-style-type: none"> ALL three correct. 		
(ii)	<p>When a small volume of HBr is added, the basic component of the buffer, CH_3COO^-, will react with (neutralise) the added H_3O^+ to produce CH_3COOH:</p> $\text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+ \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}$ <p>OR</p> $\text{CH}_3\text{COO}^- + \text{HBr} \rightarrow \text{CH}_3\text{COOH} + \text{Br}^-$ <p>Since the CH_3COOH produced is a much weaker acid than HBr, the pH does not change significantly.</p>	<ul style="list-style-type: none"> Identifies the acid / H_3O^+ ions are removed / neutralised by the buffer. 	<ul style="list-style-type: none"> Explains that the CH_3COO^- neutralises / reacts with the $\text{H}_3\text{O}^+/\text{HBr}$. 	
(iii)	$\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+$ $K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$ $1.74 \times 10^{-5} = \frac{[\text{CH}_3\text{COO}^-] \times 10^{-4.11}}{0.354}$ $[\text{CH}_3\text{COO}^-] = 0.0794 \text{ mol L}^{-1}$ $n(\text{CH}_3\text{COO}^-) = cV = 0.0794 \times 0.250$ $= 0.0198 \text{ mol}$ $m(\text{CH}_3\text{COONa}) = n \times M = 0.0198 \times 82.0$ $= 1.63 \text{ g}$	<ul style="list-style-type: none"> Correctly substitutes into K_a expression / other suitable equation. 	<ul style="list-style-type: none"> Correct process to determine mass of CH_3COONa required, with minor error. 	<ul style="list-style-type: none"> Explains how the CH_3COO^- neutralises / reacts with the H_3O^+, with supporting equation. AND Correctly calculates mass of CH_3COONa required, including correct unit and significant figures (accept 2–4).

(b)	<p>Solution A is CH_3COOH.</p> <p>CH_3COOH is a weak acid and only partially dissociates to form a low $[\text{CH}_3\text{COO}^-]$ and $[\text{H}_3\text{O}^+]$ / low [ions]. As a result, it is a poor electrical conductor. Furthermore, since it produces H_3O^+ ions, $[\text{H}_3\text{O}^+] > [\text{OH}^-]$, the pH is less than 7 as shown by yellow colour with bromothymol blue.</p> $\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+$ <p>Solution B is CH_3NH_2.</p> <p>CH_3NH_2 is a weak base and only partially dissociates to form a low $[\text{CH}_3\text{NH}_3^+]$ and $[\text{OH}^-]$ / low [ions]. As a result, it is a poor electrical conductor. Furthermore, since it produces OH^- ions, $[\text{OH}^-] > [\text{H}_3\text{O}^+]$, the pH is greater than 7 as shown by blue colour with bromothymol blue.</p> $\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$ <p>Solution C is NH_4Cl.</p> <p>NH_4Cl is an acidic salt and completely dissociates to form a high $[\text{NH}_4^+]$ and $[\text{Cl}^-]$ / high [ions]. As a result, it is a good electrical conductor. Furthermore, NH_4^+ partially dissociates in water to produce H_3O^+ ions, $[\text{H}_3\text{O}^+] > [\text{OH}^-]$, so the pH is less than 7 as shown by yellow colour, with bromothymol blue.</p> $\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$ $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$	<ul style="list-style-type: none"> Recognises that ions are required for a solution to conduct electricity. Recognises that CH_3COOH and NH_4Cl are both acidic as indicated by yellow colour in bromothymol blue OR that CH_3NH_2 is basic as indicated by blue colour in bromothymol blue. 	<ul style="list-style-type: none"> Links the electrical conductivity to the degree of dissociation and [ions] for TWO of the unknown solutions. Links the colour change in bromothymol blue of TWO of the unknown solutions to $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$. 	<ul style="list-style-type: none"> Justifies the correct identification of all THREE solutions with reference to degree of dissociation, relative [ions], including comparison of $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$, and relevant equations.
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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)(i)	$K_a = \frac{[F^-][H_3O^+]}{[HF]}$ $6.76 \times 10^{-4} = \frac{[H_3O^+]^2}{0.135}$ $[H_3O^+] = 9.55 \times 10^{-3}$ $pH = -\log 9.55 \times 10^{-3}$ $= 2.02$	<ul style="list-style-type: none"> Correctly substitutes into K_a expression. 	<ul style="list-style-type: none"> Correctly calculates pH. 	
(ii)	<p>Curve starts at pH 2.02; after 10.0 mL, the pH equals pH 3.17; curve has sharp increase in pH at the start but then flattens out.</p>  <p style="text-align: center;">Volume of NaOH added / mL</p>	<ul style="list-style-type: none"> At least ONE pH point plotted correctly AND one aspect of shape correctly drawn. 	<ul style="list-style-type: none"> Plots BOTH points correctly plus correctly drawn curve shape. 	

(b)(i)	$\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{F}^- + \text{H}_3\text{O}^+$ $K_a = \frac{[\text{F}^-][\text{H}_3\text{O}^+]}{[\text{HF}]}$ $6.76 \times 10^{-4} = \frac{9 \times [\text{H}_3\text{O}^+]}{1}$ $[\text{H}_3\text{O}^+] = \frac{6.76 \times 10^{-4}}{9}$ $= 7.51 \times 10^{-5}$ $\text{pH} = -\log 7.51 \times 10^{-5}$ $= 4.12$ <p>Since the HF : NaF ratio is within the 1:10 range, it will function as a buffer. However, $[\text{F}^-] > [\text{HF}]$; this means the buffer will be more effective at neutralising added strong acid.</p> <p>OR</p> <p>Since the pH is greater than the $\text{p}K_a$, $[\text{F}^-] > [\text{HF}]$; this means the buffer will be more effective at neutralising added strong acid.</p>	<ul style="list-style-type: none"> • Correct process to calculate pH OR Recognises solution is more effective at neutralising strong acid / pH falls in pH range of 2.17-4.17, so will function as a buffer. 	<ul style="list-style-type: none"> • Correct pH and explanation of buffer's effectiveness. 	<ul style="list-style-type: none"> • Correct pH and evaluates the effectiveness of the buffer AND justifies ONE reason why the pH rapidly increases from 18 – 22 mL.
(ii)	<p>Between 18 and 20 mL, the buffer solution is no longer effective since most of the HF has been neutralised, so the HF:NaF ratio is greater than 1:10. As a result, addition of the NaOH causes a rapid increase in the pH, as the few remaining HF are neutralised. Between 20 and 22 mL, all of the HF has been neutralised, so the added NaOH is no longer neutralised. As a result, the pH rapidly increases due to the strong base, NaOH, completely dissociating to produce a high $[\text{OH}^-]$.</p>	<ul style="list-style-type: none"> • Recognises the solution is no longer a buffer from 18–20 mL. OR Recognises the NaOH is not neutralised after 20 mL. 	<ul style="list-style-type: none"> • Explains why the pH increases from 18–20 mL because the buffer solution no longer functions effectively. OR Explains that the pH increases after 20 mL, since all the HF is neutralised, so the NaOH is in excess. 	

(c)(i)	$\text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^-$ <p>For this solution:</p> $[\text{HF}] = [\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]}$ <p>Using K_a to calculate pH:</p> $K_a = \frac{[\text{F}^-][\text{H}_3\text{O}^+]}{[\text{HF}]}$ $6.76 \times 10^{-4} = (0.135 \times \frac{25}{45}) \times \frac{[\text{H}_3\text{O}^+]^2}{10^{-14}}$ $[\text{H}_3\text{O}^+] = 9.49 \times 10^{-9} \text{ mol L}^{-1}$ <p>pH = 8.02</p>	<ul style="list-style-type: none"> Correct process to calculate pH of a basic solution. 	<ul style="list-style-type: none"> Calculates pH but does not appropriately dilute the solution. <p>OR</p>	<ul style="list-style-type: none"> Correctly calculates pH at the equivalence point AND explains why the equivalence point will be higher if HF is replaced with CH_3COOH with reference to magnitude of K_a, degree of dissociation, and $[\text{H}_3\text{O}^+]$.
(ii)	<p>Ethanoic acid has a smaller K_a than HF and is therefore a weaker acid. This means that its conjugate base, CH_3COO^-, is a stronger base than F^-. The CH_3COO^- will dissociate to a greater extent than F^-, so it will have a higher $[\text{OH}^-]$ and therefore a lower $[\text{H}_3\text{O}^+]$. As a result, the pH at the equivalence point will be higher.</p>	<ul style="list-style-type: none"> Recognises CH_3COOH is a weaker acid than HF. 	<p>Links pH at equivalence point to magnitude of K_a and degree of dissociation of the ethanoate ion / $[\text{OH}^-]$.</p>	

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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 – 17	08 – 13	14 – 18	19 – 24