

Assessment Schedule – 2023**Chemistry: Demonstrate understanding of equilibrium principles in aqueous systems (91392)****Evidence Statement**

Q	Evidence	Achievement	Merit	Excellence
ONE (a)(i)	$\text{Pb}(\text{OH})_2 \rightleftharpoons \text{Pb}^{2+} + 2\text{OH}^-$	<ul style="list-style-type: none"> Correct equilibrium equation and K_s expression. 		
(ii)	$K_s = [\text{Pb}^{2+}][\text{OH}^-]^2$			
(iii)	$[\text{H}_3\text{O}^+] = 10^{-9.55}$ $[\text{OH}^-] = \frac{1 \times 10^{-14}}{10^{-9.55}} = 3.55 \times 10^{-5} \text{ mol L}^{-1}$ $8 \times 10^{-17} = [\text{Pb}^{2+}] \times (3.55 \times 10^{-5})^2$ $[\text{Pb}^{2+}] = 6.35 \times 10^{-8} \text{ mol L}^{-1} = \text{solubility}$	<ul style="list-style-type: none"> Substitutes into K_s expression. OR Correct $[\text{OH}^-]$.	<ul style="list-style-type: none"> Correct process for determining solubility with one minor error. 	
(iv)	When the pH is decreased, $[\text{H}_3\text{O}^+]$ increases. The H_3O^+ removes OH^- from the equilibrium and neutralises it according to the following equation: $\text{H}_3\text{O}^+ + \text{OH}^- \rightarrow 2\text{H}_2\text{O}$ Since $[\text{OH}^-]$ decreases, the equilibrium will speed up the rate of the forward reaction to replace some of the OH^- . As a result, more solid $\text{Pb}(\text{OH})_2$ will dissolve, so solubility increases.	<ul style="list-style-type: none"> Recognises that a decrease in pH will increase $[\text{H}_3\text{O}^+]$. 	<ul style="list-style-type: none"> Explains that a decrease in $[\text{OH}^-]$ due to reaction with H_3O^+ will favour formation of more OH^- in solution / favour forward reaction. 	<ul style="list-style-type: none"> Correct solubility, including units and 2-4 significant figures. AND Fully explains, using equilibrium principles, the effect of decreasing the pH on the solubility of $\text{Pb}(\text{OH})_2$ (replacing the removed hydroxide ions idea), including equation.

(b)(i)	$IP = [Pb^{2+}][I^-]^2$ $= \left(\frac{25}{60} \times 0.00741\right) \times \left(\frac{35}{60} \times 0.00613\right)^2$ $= 0.00309 \times (0.00358)^2$ $= 3.95 \times 10^{-8}$ <p>Since $IP > K_s$, a precipitate of PbI_2 will form.</p>	<ul style="list-style-type: none"> • EITHER correct $[Pb^{2+}]$ OR $[I^-]$ • Correct comparison of IP with K_s. 	<ul style="list-style-type: none"> • Correct process to calculate IP and compare with K_s (may have one dilution incorrect). 	<ul style="list-style-type: none"> • Correct calculation (2-4 significant figures) and comparison with K_s to show that PbI_2 will form a precipitate. <p>AND</p> <p>Uses equilibrium principles to fully explain the effect of adding a common ion on the solubility of PbI_2 (using up iodide ions idea) OR explains with reference to increase in IP.</p>
(ii)	<p>NaI circled. The addition of NaI would increase the concentration of the common ion, I^-. This would cause an increase in the rate of the reverse reaction to use up some of the I^-. As a result, more solid PbI_2 would form, so the solubility of PbI_2 would decrease.</p>	<ul style="list-style-type: none"> • Circles NaI and recognises I^- is a common ion / $[I^-]$ increases. 	<ul style="list-style-type: none"> • Explains the addition of/increase in I^- favours the reverse reaction. 	

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)	$[\text{CO}_3^{2-}] = \frac{50}{85} \times 0.211 = 0.124 \text{ mol L}^{-1}$ $[\text{HCO}_3^-] = \frac{35}{85} \times 0.861 = 0.355 \text{ mol L}^{-1}$ $\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{CO}_3^{2-} + \text{H}_3\text{O}^+$ $K_a = \frac{[\text{CO}_3^{2-}][\text{H}_3\text{O}^+]}{[\text{HCO}_3^-]}$ $10^{-10.3} = \frac{0.124 \times [\text{H}_3\text{O}^+]}{0.355}$ $[\text{H}_3\text{O}^+] = 1.43 \times 10^{-10}, \text{pH} = 9.84$	<ul style="list-style-type: none"> Calculates $[\text{CO}_3^{2-}]$ or $[\text{HCO}_3^-]$ OR Substitutes correctly into K_a expression or the Henderson-Hasselbalch equation.	<ul style="list-style-type: none"> Correct process for calculating pH of buffer solution with one minor error, e.g., error in one of the dilutions. 	<ul style="list-style-type: none"> Correctly calculates pH of buffer solution (2-4 significant figures).
(ii)	The solution will act as an effective buffer over a pH range 9.3–11.3 ($\text{p}K_a \pm 1$). The solution made by the student has a pH of 8.50 which does not fall within this pH range.	<ul style="list-style-type: none"> Recognises why the solution will not act effectively as a buffer. 		AND
(iii)	For the solution to function as a buffer equally effective at neutralising added strong acid or base, the ratio of $\text{HCO}_3^-:\text{CO}_3^{2-}$ needs to be a 1:1 ratio. At this ratio, the pH of the buffer solution will be equal to the $\text{p}K_a$ of 10.3. Since the pH of 8.50 is lower than 10.3, some strong base (e.g. NaOH) needs to be added until the pH reaches 10.3.	<ul style="list-style-type: none"> Identifies that carbonate ions / (strong) base needs to be added. OR Identifies pH of 10.3 or ratio of 1:1 required.	<ul style="list-style-type: none"> Explains how the solution can be altered by adding more base with reference to either the ratio of $\text{HCO}_3^-:\text{CO}_3^{2-}$ OR pH 10.3. 	<ul style="list-style-type: none"> Elaborates on how the solution needs to be altered to produce a buffer equally effective at neutralising both added acid or base. Must mention adding carbonate ions or a strong base like NaOH to increase pH. Need to refer to BOTH 1:1 ratio and pH of 10.3 ($\text{p}K_a$).

<p>(b)(i)</p> <p>(ii)</p>	<p>$\text{Cl}^- > \text{NH}_4^+ > \text{NH}_3 = \text{H}_3\text{O}^+ > \text{OH}^-$</p> <p>$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$</p> $K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$ $5.75 \times 10^{-10} = \frac{(10^{-5.25})^2}{[\text{NH}_4^+]}$ <p>$[\text{NH}_4^+] = 0.0550 \text{ mol L}^{-1}$</p> <p>$n(\text{NH}_4^+) = 0.0550 \times 0.500 = 0.0275 \text{ mol}$</p> <p>$m(\text{NH}_4\text{Cl}) = 0.0275 \times 53.5 = 1.47 \text{ g}$</p>	<ul style="list-style-type: none"> All three boxes correct. Correct $[\text{NH}_4^+]$. 	<ul style="list-style-type: none"> Correct mass of NH_4Cl, including unit and significant figures (2 – 4). 	
<p>(iii)</p>	<p>NH_4Cl is an acidic salt and completely dissociates to form high $[\text{NH}_4^+]$ and $[\text{Cl}^-]$. As a result, NH_4Cl has a high [ions] and is therefore a good electrical conductor.</p> <p>$\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$</p> <p>The NH_4^+ ion is only a weak acid and therefore partially dissociates to form a relatively low $[\text{H}_3\text{O}^+]$. (Since $[\text{H}_3\text{O}^+] > [\text{OH}^-]$, the pH is still less than 7, but not as low as for a strong acid.)</p> <p>$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$</p> <p>$\text{HBr}$ is a strong acid and completely dissociates to form high $[\text{Br}^-]$ and $[\text{H}_3\text{O}^+]$.</p> <p>$\text{HBr} + \text{H}_2\text{O} \rightarrow \text{Br}^- + \text{H}_3\text{O}^+$</p> <p>As a result, HBr has a high [ions], and is therefore a good electrical conductor. Furthermore, since HBr completely dissociates, it produces a higher $[\text{H}_3\text{O}^+]$ than NH_4^+, and therefore has a lower pH.</p>	<ul style="list-style-type: none"> Recognises a solution requires mobile ions for electrical conductivity. <p>OR</p> <p>pH is dependent on $[\text{H}_3\text{O}^+]$.</p>	<ul style="list-style-type: none"> Links electrical conductivity to degree of dissociation and high [ions] for BOTH solutions. Links pH to the degree of dissociation and $[\text{H}_3\text{O}^+]$ for BOTH solutions. 	<ul style="list-style-type: none"> Fully compares the pH and electrical conductivity of BOTH solutions, including relevant equations (at least two correct as shown in schedule).

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

(iii)	A suitable indicator would need to change colour in the vertical section of the curve which stretches from a pH of approximately 6 – 10. The pK_a of the indicator would ideally need to be close to the equivalence point pH, since the pH range over which the indicator will change colour is equal to the $pK_a \pm 1$. This would mean the indicator would definitely change colour in the pH range 6 – 10.	<ul style="list-style-type: none"> Recognises the indicator needs to change colour in the vertical section of the curve / close to the equivalence point. 	<ul style="list-style-type: none"> Explains the pH range over which a suitable indicator would need to change colour and relates this to its pK_a. 	
(c)(i)	Between 8 – 12 mL, a buffer solution consisting $HCOOH$ and $HCOO^-$ is present in the conical flask. When the KOH is added, the OH^- ions are neutralised by the buffer according to the equation below: $HCOOH + OH^- \rightarrow HCOO^- + H_2O$ The strong base, KOH , is reacted to form a weak base, $HCOO^-$, which has less effect on the pH. There is a small change in $[HCOOH]$: $[HCOO^-]$, so only a small change in pH.	<ul style="list-style-type: none"> Recognises a buffer solution is present between 8 and 12 mL. 	<ul style="list-style-type: none"> Explains how the $HCOOH$ present in the buffer solution neutralises/reacts with/removes the added KOH. 	<ul style="list-style-type: none"> Fully explains why the pH does not significantly change between 8 – 12 mL, including relevant equation (allow minor error in equation for E7).
(ii)	After 23 mL of KOH has been added: $n(KOH) = cv = 0.0893 \times 0.003 = 2.679 \times 10^{-4} \text{ mol}$ $c(KOH) = \frac{n}{V} = \frac{2.679 \times 10^{-4}}{0.038}$ $= 7.05 \times 10^{-3} \text{ mol L}^{-1}$ $[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1 \times 10^{-14}}{7.05 \times 10^{-3}}$ $= 1.42 \times 10^{-12}$ $pH = -\log 1.42 \times 10^{-12} = 11.8$	<ul style="list-style-type: none"> Correctly calculates $n(KOH)$ remaining in excess. OR Shows correct process to calculate diluted $[KOH]$.	<ul style="list-style-type: none"> Correctly calculates diluted $c(KOH)$. 	AND Correctly calculates pH (2-4 significant figures) after 23 mL has been added.

N0	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 – 8	9 – 13	14 – 18	19 – 24