

## 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"> <li>Correct equilibrium equation and <math>K_s</math> expression.</li> </ul>		
(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"> <li>Substitutes into <math>K_s</math> expression.</li> </ul> OR Correct $[\text{OH}^-]$ .	<ul style="list-style-type: none"> <li>Correct process for determining solubility with one minor error.</li> </ul>	
(iv)	When the pH is decreased, $[\text{H}_3\text{O}^+]$ increases. The $\text{H}_3\text{O}^+$ removes $\text{OH}^-$ 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 $\text{OH}^-$ . As a result, more solid $\text{Pb}(\text{OH})_2$ will dissolve, so solubility increases.	<ul style="list-style-type: none"> <li>Recognises that a decrease in pH will increase <math>[\text{H}_3\text{O}^+]</math>.</li> </ul>	<ul style="list-style-type: none"> <li>Explains that a decrease in <math>[\text{OH}^-]</math> due to reaction with <math>\text{H}_3\text{O}^+</math> will favour formation of more <math>\text{OH}^-</math> in solution / favour forward reaction.</li> </ul>	<ul style="list-style-type: none"> <li>Correct solubility, including units and 2-4 significant figures.</li> </ul> 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 <math>IP &gt; K_s</math>, a precipitate of <math>PbI_2</math> will form.</p>	<ul style="list-style-type: none"> <li>• EITHER correct <math>[Pb^{2+}]</math> OR <math>[I^-]</math></li> <li>• Correct comparison of IP with <math>K_s</math>.</li> </ul>	<ul style="list-style-type: none"> <li>• Correct process to calculate IP and compare with <math>K_s</math> (may have one dilution incorrect).</li> </ul>	<ul style="list-style-type: none"> <li>• Correct calculation (2-4 significant figures) and comparison with <math>K_s</math> to show that <math>PbI_2</math> will form a precipitate.</li> </ul> <p>AND</p> <p>Uses equilibrium principles to fully explain the effect of adding a common ion on the solubility of <math>PbI_2</math> (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, <math>I^-</math>. This would cause an increase in the rate of the reverse reaction to use up some of the <math>I^-</math>. As a result, more solid <math>PbI_2</math> would form, so the solubility of <math>PbI_2</math> would decrease.</p>	<ul style="list-style-type: none"> <li>• Circles NaI and recognises <math>I^-</math> is a common ion / <math>[I^-]</math> increases.</li> </ul>	<ul style="list-style-type: none"> <li>• Explains the addition of/increase in <math>I^-</math> favours the reverse reaction.</li> </ul>	

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

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

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)	$\text{HCOOH} + \text{H}_2\text{O} \rightleftharpoons \text{HCOO}^- + \text{H}_3\text{O}^+$ $K_a = \frac{[\text{HCOO}^-][\text{H}_3\text{O}^+]}{[\text{HCOOH}]}$ $1.82 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+]^2}{0.119}$ $[\text{H}_3\text{O}^+] = 4.65 \times 10^{-3} \text{ mol L}^{-1}$ $\text{pH} = -\log 4.65 \times 10^{-3} = 2.33$	<ul style="list-style-type: none"> <li>Correct process for calculating pH of HCOOH solution (correct substitution).</li> </ul>	<ul style="list-style-type: none"> <li>Correctly calculates pH of HCOOH solution (2 – 4 significant figures).</li> </ul>	
(b)(i)  (ii)	<p>Approximate pH range: 8.0–8.4</p> <p>For this solution:  <math display="block">\text{HCOO}^- + \text{H}_2\text{O} \rightleftharpoons \text{HCOOH} + \text{OH}^-</math></p> <p>For this solution:</p> $[\text{HCOOH}] = [\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]}$ <p>Using <math>K_a</math> to calculate pH:</p> $K_a = \frac{[\text{HCOO}^-][\text{H}_3\text{O}^+]}{[\text{HCOOH}]}$ $1.82 \times 10^{-4} = \frac{\left(0.119 \times \frac{15}{35}\right) \times [\text{H}_3\text{O}^+]^2}{10^{-14}}$ $[\text{H}_3\text{O}^+] = 5.97 \times 10^{-9} \text{ mol L}^{-1}$ $\text{pH} = 8.22$	<ul style="list-style-type: none"> <li>Estimate of pH falls within expected range (don't accept a range from candidates).</li> <li>Correct process to calculate pH (cannot use [KOH] in calculation). Looking for correct substitution.</li> </ul>	<ul style="list-style-type: none"> <li>Calculates pH at the equivalence point (allow error in dilution; pH of 8.41 if uses undiluted concentration).</li> </ul>	<ul style="list-style-type: none"> <li>Calculates the pH at the equivalence point (accepts 2-4 significant figures).            AND            Explains how to choose a suitable indicator (should relate to the approximate <math>pK_a</math> of the indicator).</li> </ul>

(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"> <li>Recognises the indicator needs to change colour in the vertical section of the curve / close to the equivalence point.</li> </ul>	<ul style="list-style-type: none"> <li>Explains the pH range over which a suitable indicator would need to change colour and relates this to its <math>pK_a</math>.</li> </ul>	
(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"> <li>Recognises a buffer solution is present between 8 and 12 mL.</li> </ul>	<ul style="list-style-type: none"> <li>Explains how the <math>HCOOH</math> present in the buffer solution neutralises/reacts with/removes the added <math>KOH</math>.</li> </ul>	<ul style="list-style-type: none"> <li>Fully explains why the pH does not significantly change between 8 – 12 mL, including relevant equation (allow minor error in equation for E7).</li> </ul>
(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"> <li>Correctly calculates <math>n(KOH)</math> remaining in excess.</li> </ul> OR Shows correct process to calculate diluted $[KOH]$ .	<ul style="list-style-type: none"> <li>Correctly calculates diluted <math>c(KOH)</math>.</li> </ul>	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