## Assessment Schedule - 2019

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

## Evidence Statement

| Q | Evidence | Achievement | Merit | Excellence |
| :---: | :---: | :---: | :---: | :---: |
| ONE <br> (a)(i) <br> (ii) <br> (iii) <br> (iv) | $\mathrm{Zn}(\mathrm{OH})_{2} \rightleftharpoons \mathrm{Zn}^{2+}+2 \mathrm{OH}^{-}$ $K_{\mathrm{s}}=\left[\mathrm{Zn}^{2+}\right]\left[\mathrm{OH}^{-}\right]^{2}$ <br> Let solubility be ' $s$ ': $\begin{aligned} & {\left[\mathrm{Zn}^{2+}\right]=s} \\ & {\left[\mathrm{OH}^{-}\right]=2 s} \end{aligned}$ $K_{\mathrm{s}}=4 s^{3}$ $4 s^{3}=3.80 \times 10^{-17}$ $s=2.12 \times 10^{-6} \mathrm{~mol} \mathrm{~L}^{-1}$ <br> $\left[\mathrm{Zn}^{2+}\right]=2.12 \times 10^{-6} \mathrm{~mol} \mathrm{~L}^{-1}$ <br> $\left[\mathrm{OH}^{-}\right]=4.24 \times 10^{-6} \mathrm{~mol} \mathrm{~L}^{-1}$ <br> $K_{\mathrm{s}}=\left[\mathrm{Zn}^{2+}\right]\left[\mathrm{OH}^{-}\right]^{2}$ $3.80 \times 10^{-17}=\left(\frac{0.210}{2}\right) \times\left[\mathrm{OH}^{-}\right]^{2}$ $\left[\mathrm{OH}^{-}\right]=1.90 \times 10^{-8} \mathrm{~mol} \mathrm{~L}^{-1}$ <br> Note: $\frac{0.210}{2}$ due to dilution factor from adding two solutions together. | Correct equilibrium equation and $K_{\mathrm{s}}$ expression. <br> Method correct for determining solubility. <br> Recognises $\mathrm{Zn}^{2+}$ is the common ion. OR <br> ONE step correct. <br> E.g. correct substitution into $K_{\mathrm{s}}$ excluding dilution / 0.105. | Correct solubility of $\mathrm{Zn}(\mathrm{OH})_{2}$ calculated, including $\left[\mathrm{Zn}^{2+}\right]$ and $\left[\mathrm{OH}^{-}\right]$. <br> Calculates $\left[\mathrm{OH}^{-}\right]$, but does not take into account the dilution factor. | Calculates correct $\left[\mathrm{OH}^{-}\right]$. <br> Including sig fig and units |
| (b) | $\mathrm{Zn}(\mathrm{OH})_{2} \rightleftharpoons \mathrm{Zn}^{2+}+2 \mathrm{OH}^{-}$ <br> When the $\mathrm{OH}^{-}$ions are in excess, the $\mathrm{Zn}^{2+}$ ions can form a complex ion: $\mathrm{Zn}^{2+}+4 \mathrm{OH}^{-} \rightarrow\left[\mathrm{Zn}(\mathrm{OH})_{4}\right]^{2-}$ <br> A decrease in $\left[\mathrm{Zn}^{2+}\right]$ will result in the forward reaction being favoured, to restore equilibrium / minimise the change. <br> This causes more solid $\mathrm{Zn}(\mathrm{OH})_{2}$ to dissolve / the solubility of $\mathrm{Zn}(\mathrm{OH})_{2}$ increases / so that $\left[\mathrm{Zn}^{2+}\right]\left[\mathrm{OH}^{-}\right]$will again equal $K_{\mathrm{s}}$. | Recognises solubility increases due to formation of complex ion. | Explains that formation of a complex ion will decrease $\left[\mathrm{Zn}^{2+}\right]$. | Fully explains, using equilibrium principles, how the solubility of the $\mathrm{Zn}(\mathrm{OH})_{2}$ increases when excess NaOH is added. Must include complex ion equation. |

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| (c) | $\begin{aligned} & {\left[\mathrm{Zn}^{2+}\right]=\frac{20}{50} \times 0.0242=9.68 \times 10^{-3} \mathrm{~mol} \mathrm{~L}^{-1}} \\ & {\left[\mathrm{OH}^{-}\right]=\frac{30}{50} \times 1 \times 10^{-14} / 10^{-13.1}=0.0755 \mathrm{~mol} \mathrm{~L}^{-1}} \\ & \mathrm{IP}=\left[\mathrm{Zn}^{2+}\right]\left[\mathrm{OH}^{-}\right]^{2}=9.68 \times 10^{-3} \times(0.0755)^{2} \\ & \quad=5.52 \times 10^{-5} \quad\left(5.53 \times 10^{-5}\right) \end{aligned}$ <br> Since IP $>K_{\mathrm{s}}$, a precipitate of $\mathrm{Zn}(\mathrm{OH})_{2}$ will form. | Correct substitution into $Q_{\mathrm{s}}$ (IP) expression. <br> OR <br> Correct $\left[\mathrm{Zn}^{2+}\right]$ or $\left[\mathrm{OH}^{-}\right]$. | Correct process to determine $Q_{\mathrm{s}}$ and compare with $K_{\mathrm{s}}$. | Correct calculation and comparison with $K_{\mathrm{s}}$ to determine whether $\mathrm{Zn}(\mathrm{OH})_{2}$ will form a precipitate. |
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| NØ | N1 | $\mathbf{N 2}$ | A3 | A4 | M5 | M6 | E7 | E8 |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| No response; <br> no relevant evidence. | 1 a | 2 a | 3 a | 4 a | 2 m | 3 m | 1 e |  |


| Q | Evidence | Achievement | Merit | Excellence |
| :---: | :---: | :---: | :---: | :---: |
| TWO <br> (a)(i) <br> (ii) | $\mathrm{Na}^{+}, \mathrm{HCOO}^{-}, \mathrm{HCOOH}, \mathrm{OH}^{-}, \mathrm{H}_{3} \mathrm{O}^{+}$ <br> After 12.5 mL NaOH has been added, it is halfway to the equivalence point. <br> This means that $[\mathrm{HCOOH}]=\left[\mathrm{HCOO}^{-}\right]$. <br> Therefore pH equals $\mathrm{p} K_{\mathrm{a}}$. | THREE species identified. $\begin{aligned} & {[\mathrm{HCOOH}]=\left[\mathrm{HCOO}^{-}\right]} \\ & \mathrm{OR} \\ & \mathrm{pH}=\mathrm{p} K_{\mathrm{a}} \end{aligned}$ | $\begin{aligned} & {[\mathrm{HCOOH}]=\left[\mathrm{HCOO}^{-}\right]} \\ & \text {Therefore } \mathrm{pH}=\mathrm{p} K_{\mathrm{a}} \text {. } \end{aligned}$ |  |
| (b)(i) <br> (ii) | Tick cresol red. <br> Indicators change colour at a $\mathrm{pH} \pm 1$ of the $\mathrm{pK}_{\mathrm{a}} /$ near the $\mathrm{pK}_{\mathrm{a}}$. <br> Therefore, cresol red should be used as it will change near the equivalence point/steepest part of the curve, whereas thymol blue and bromocresol green will change before the equivalence point/steepest part of the curve. $\begin{aligned} & {\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]=\sqrt{\frac{K_{\mathrm{a}} \times K_{\mathrm{w}}}{\left[\mathrm{CH}_{3} \mathrm{CH}_{2} \mathrm{NH}_{2}\right]}}} \\ & {\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]=\sqrt{\frac{1.82 \times 10^{-4} \times 1 \times 10^{-14}}{0.0778}}} \\ & {\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]=4.84 \times 10^{-9} \mathrm{~mol} \mathrm{~L}^{-1}} \\ & \mathrm{pH}=-\log \left[\mathrm{H}_{3} \mathrm{O}^{+}\right]=8.32 \end{aligned}$ <br> OR <br> For this solution, $\mathrm{HCOO}^{-}+\mathrm{H}_{2} \mathrm{O} \rightleftharpoons \mathrm{HCOOH}+\mathrm{OH}^{-}$ | Identifies cresol red. <br> OR <br> Recognises an indicator is chosen to change colour over the vertical section of the curve / at the equivalence point. <br> Correct process for determining pH at equivalence point. <br> OR <br> One correct step. | Explanation with reference to the $\mathrm{p} K_{\mathrm{a}}$ of why cresol red is the appropriate indicator whereas the other two indicators are not. <br> pH calculated at equivalence point with incorrect dilution. | Full explanation for indicator choice . <br> AND <br> pH at equivalence point. |
| (c) | After 28 mL NaOH added: $\begin{aligned} & n(\text { unreacted } \mathrm{NaOH})=c V=0.140 \times 0.003=4.2 \times 10^{-4} \mathrm{~mol} \\ & c(\mathrm{NaOH})=\frac{n}{V}=\frac{4.2 \times 10^{-4}}{0.048}=8.75 \times 10^{-3} \mathrm{~mol} \mathrm{~L}^{-1} \\ & {\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]=\frac{1 \times 10^{-14}}{8.75 \times 10^{-3}}=1.14 \times 10^{-12} \mathrm{~mol} \mathrm{~L}^{-1}} \\ & \mathrm{pH}=-\log 1.14 \times 10^{-12}=11.9 \end{aligned}$ | Correct $n(\mathrm{NaOH})$. <br> OR <br> One correct step. | Correct process but one error in calculation. | Correct pH. |

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| NØ | $\mathbf{N 1}$ | $\mathbf{N 2}$ | $\mathbf{A 3}$ | A4 | M5 | M6 | E7 | E8 |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| No response; <br> no relevant evidence. | 1 a | 2 a | 3 a | 4 a | 2 m | 3 m | 1 e |  |

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| Q | Evidence | Achievement | Merit | Excellence |
| :---: | :---: | :---: | :---: | :---: |
| THREE <br> (a)(i) | $\mathrm{CH}_{3} \mathrm{COOH}$ has the smaller $\mathrm{p} K_{\mathrm{a}} /$ larger $K_{\mathrm{a}}$, so it will be a stronger acid than $\mathrm{NH}_{4}{ }^{+}$. <br> This means $\mathrm{CH}_{3} \mathrm{COOH}$ will dissociate to a greater extent to produce a higher $\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]$than $\mathrm{NH}_{4}{ }^{+}$, so $\mathrm{CH}_{3} \mathrm{COOH}$ will have a lower pH . | Recognises $\mathrm{p} K_{\mathrm{a}}$ is a measure of the extent of dissociation of a weak acid. | Links magnitude of $\mathrm{p} K_{\mathrm{a}}$ (or $K_{\mathrm{a}}$ ) to degree of dissociation, $\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]$and pH . <br> OR | Full explanation of the pH and electrical conductivity of BOTH solutions, i.e. links degree of dissociation to BOTH $\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]$and pH AND [ions] and electrical |
| (ii) | The electrical conductivity of a solution depends upon [ions]. $\mathrm{NH}_{4} \mathrm{Cl}$ is an acidic salt, so it completely dissociates / dissolves to produce a relatively high $\left[\mathrm{NH}_{4}{ }^{+}\right]$and $\left[\mathrm{Cl}^{-}\right]$. This makes $\mathrm{NH}_{4} \mathrm{Cl}$ a good electrical conductor. | Recognises electrical conductivity in a solution requires ions. | Links the [ions] to the electrical conductivity of BOTH solutions. | conductivity, including supporting equations. |
|  | $\mathrm{NH}_{4} \mathrm{Cl} \rightarrow \mathrm{NH}_{4}^{+}+\mathrm{Cl}^{-}$ |  |  |  |
|  | However, $\mathrm{CH}_{3} \mathrm{COOH}$ is a weak acid and only partially dissociates to produce a relatively low $\left[\mathrm{CH}_{3} \mathrm{COO}^{-}\right]$and $\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]$. This makes $\mathrm{CH}_{3} \mathrm{COOH}$ a poorer electrical conductor than $\mathrm{NH}_{4} \mathrm{Cl}$. $\mathrm{CH}_{3} \mathrm{COOH}+\mathrm{H}_{2} \mathrm{O} \rightleftharpoons \mathrm{CH}_{3} \mathrm{COO}^{-}+\mathrm{H}_{3} \mathrm{O}^{+}$ |  |  |  |
| (iii) | $\mathrm{CH}_{3} \mathrm{COOH}+\mathrm{H}_{2} \mathrm{O} \rightleftharpoons \mathrm{CH}_{3} \mathrm{COO}^{-}+\mathrm{H}_{3} \mathrm{O}^{+}$ |  |  |  |
|  | $K_{\mathrm{a}}=\frac{\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]\left[\mathrm{CH}_{3} \mathrm{COO}^{-}\right]}{\left[\mathrm{CH}_{3} \mathrm{COOH}\right]}$ | Correct process for determining [ $\left.\mathrm{CH}_{3} \mathrm{COOH}\right]$. (correct substitution |  |  |
|  | $10^{-4.76}=\frac{\left(1.78 \times 10^{-3}\right)^{2}}{\Gamma}$ | into formula) <br> OR |  |  |
|  | $10=\left[\mathrm{CH}_{3} \mathrm{COOH}\right]$ |  | Correct [ $\left.\mathrm{CH}_{3} \mathrm{COOH}\right]$. |  |
|  | $\left[\mathrm{CH}_{3} \mathrm{COOH}\right]=0.182 \mathrm{~mol} \mathrm{~L}^{-1}$ |  |  |  |


| (b)(i) | $\begin{aligned} & \mathrm{CH}_{3} \mathrm{COOH}+\mathrm{H}_{2} \mathrm{O} \rightleftharpoons \mathrm{CH}_{3} \mathrm{COO}^{-}+\mathrm{H}_{3} \mathrm{O}^{+} \\ & K_{\mathrm{a}}=\frac{\left[\mathrm{CH}_{3} \mathrm{COO}^{-}\right]\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]}{\left[\mathrm{CH}_{3} \mathrm{COOH}\right]} \\ & 10^{-4.76}=2 \times \frac{\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]}{5} \\ & {\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]=4.34 \times 10^{-5} \mathrm{~mol} \mathrm{~L}^{-1}} \\ & \mathrm{pH}=-\log 4.34 \times 10^{-5}=4.36 \end{aligned}$ | Correct process for determining pH . OR <br> One correct step. | Correct pH. |  |
| :---: | :---: | :---: | :---: | :---: |
| (ii) | $\left[\mathrm{CH}_{3} \mathrm{COOH}\right]>\left[\mathrm{CH}_{3} \mathrm{COO}^{-}\right] / \mathrm{pH}<\mathrm{pK}_{\mathrm{a}}$ <br> Therefore the buffer solution is more effective at neutralising strong base: $\mathrm{CH}_{3} \mathrm{COOH}+\mathrm{OH}^{-} \rightarrow \mathrm{CH}_{3} \mathrm{COO}^{-}+\mathrm{H}_{2} \mathrm{O}$ | Equation. <br> OR $\left[\mathrm{CH}_{3} \mathrm{COOH}\right]>\left[\mathrm{CH}_{3} \mathrm{COO}^{-}\right]$ | Links ratio of $\left[\mathrm{CH}_{3} \mathrm{COOH}\right]:\left[\mathrm{CH}_{3} \mathrm{COO}^{-}\right]$ to buffer effectiveness, including equation. | Full explanation of buffer behaviour for (ii) and (iii). |
| (iii) | When water is added, the ratio of $\mathrm{CH}_{3} \mathrm{COOH}$ to $\mathrm{CH}_{3} \mathrm{COO}^{-}$is unchanged, so the pH of the buffer solution is unaffected. | Recognises pH remains unchanged. | Explains effect of dilution on pH in terms of ratio between $\mathrm{CH}_{3} \mathrm{COOH}$ and $\mathrm{CH}_{3} \mathrm{COO}^{-}$. |  |


| NØ | N1 | N2 | A3 | A4 | M5 | M6 | E7 | E8 |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| No response; no relevant evidence. | 1a | 2a | 3a | 4 a | 2m | 3 m | 1 e | 2 e |

## Cut Scores

| Not Achieved | Achievement | Achievement with Merit | Achievement with Excellence |
| :---: | :---: | :---: | :---: |
| $0-6$ | $7-13$ | $14-19$ | $20-24$ |

