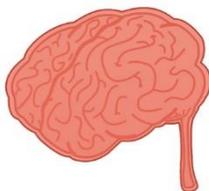


SUGGESTED ANSWERS – please email any corrections / suggestions



## Level 3 Chemistry

### 91392 Demonstrate understanding of equilibrium principles in aqueous systems

Credits: Five

Achievement	Achievement with Merit	Achievement with Excellence
Demonstrate understanding of equilibrium principles in aqueous systems	Demonstrate in-depth understanding of equilibrium principles in aqueous systems	Demonstrate comprehensive understanding of equilibrium principles in aqueous systems

You should attempt ALL the questions in this booklet.

A periodic table is provided in the Resource Sheet.

If you need more room for any answer, use the extra space provided at the back of this booklet and clearly number the question.

Check that this booklet has pages 2–9 in the correct order and that none of these pages is blank.

**YOU MUST HAND THIS BOOKLET TO THE SUPERVISOR AT THE END OF THE EXAMINATION.**

TOTAL

ASSESSOR'S USE ONLY

**QUESTION ONE**

Strontium carbonate,  $\text{SrCO}_3$ , is a sparingly soluble salt. It is widely used in the ceramics industry as an ingredient in glazes.

$K_s(\text{SrCO}_3) = 1.10 \times 10^{-10}$  at  $25^\circ\text{C}$ .  $M(\text{SrCO}_3) = 148 \text{ g mol}^{-1}$ .

- (a) Write the solubility product expression,  $K_s$ , for strontium carbonate.

$$K_s = [\text{Sr}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})]$$


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- (b) (i) Calculate the solubility of strontium carbonate,  $\text{SrCO}_3$ , at this temperature.

$$\text{AB type} \quad s = \sqrt{K_s} = 1.05 \times 10^{-5} \text{ mol L}^{-1}$$


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- (ii) Calculate the mass of strontium carbonate that will dissolve in 150 mL of water to make a saturated solution at  $25^\circ\text{C}$ .

$$s = 1.05 \times 10^{-5} \text{ mol L}^{-1} = 1.05 \times 10^{-5} \times 148 = 1.55 \times 10^{-3} \text{ g L}^{-1}$$


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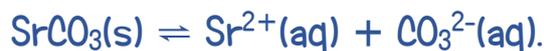
$$\text{Therefore in 150 mL. } 1.55 \times 10^{-3} \times 150/1000 = 2.33 \times 10^{-4} \text{ g (3 s.f.)}$$


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- (c) Explain how the solubility of strontium carbonate,  $\text{SrCO}_3$ , will change if added to 150 mL of a  $1.00 \text{ mol L}^{-1}$  hydrochloric acid solution, HCl. Support your answer with balanced equations.

*No calculations are necessary.*

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The acid will react with the carbonate ion  $2\text{H}_3\text{O}^+ + \text{CO}_3^{2-} \rightarrow 3\text{H}_2\text{O} + \text{CO}_2$  and  $\text{CO}_2$  gas will be formed

Removal of  $\text{CO}_2$  means the concentration of  $\text{CO}_3^{2-}(\text{aq})$  decreases and the equilibrium position will shift to the right (to replace it) and the solubility will increase.

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- (d) A maximum of 0.600 g  $\text{Pb}(\text{NO}_3)_2$  can be added to 1.50 L of sodium bromide solution,  $\text{NaBr}(\text{aq})$  without forming a precipitate of lead bromide,  $\text{PbBr}_2$ .

$$K_s (\text{PbBr}_2) = 6.60 \times 10^{-6} \text{ at } 25^\circ\text{C}. \quad M(\text{Pb}(\text{NO}_3)_2) = 331 \text{ g mol L}^{-1}.$$

Calculate the concentration of the sodium bromide solution.

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When  $Q$  or  $IP = K_s$  then the solution will be saturated but a precipitate will not have formed.

$$Q = [\text{Pb}^{2+}(\text{aq})] [\text{Br}^{-}(\text{aq})]^2$$

$n(\text{Pb}(\text{NO}_3)_2) = m/M = 0.600/331 = 1.81 \times 10^{-3}$  mol; this is added to 1.5 L so the concentration of  $\text{Pb}^{2+}(\text{aq}) = n/v = 1.81 \times 10^{-3}/1.5 = 1.21 \times 10^{-3} \text{ mol L}^{-1}$

$$6.60 \times 10^{-6} = 1.21 \times 10^{-3} \times [\text{Br}^{-}(\text{aq})]^2$$

Therefore  $[\text{Br}^{-}(\text{aq})]^2 = 5.46 \times 10^{-3}$  and  $[\text{Br}^{-}(\text{aq})] = 0.0739 \text{ mol L}^{-1}$

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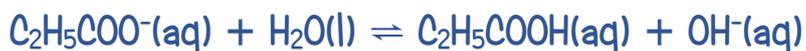
**Question Two**

(a) Sodium propanoate,  $\text{C}_2\text{H}_5\text{COONa}$  dissolves in water to form a weakly basic solution.  $\text{p}K_{\text{a}}$  ( $\text{C}_2\text{H}_5\text{COOH}$ ) is 4.87.

(i) Write an equation for when sodium propanoate dissolves in water.



(ii) Write an equation to show the reaction occurring in an aqueous solution of  $\text{C}_2\text{H}_5\text{COONa}$ .



(iii) List all the species present in an aqueous solution of  $\text{C}_2\text{H}_5\text{COONa}$ , in order of decreasing concentration. Do not include water.



(b) Calculate the pH of  $0.0175 \text{ mol L}^{-1}$   $\text{C}_2\text{H}_5\text{COONa}$  solution.

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$$K_{\text{b}} = [\text{OH}^{\text{-}}]^2 / [\text{salt}]$$

$$\text{p}K_{\text{a}} (\text{C}_2\text{H}_5\text{COOH}) \text{ is } 4.87$$

$$\text{so } \text{p}K_{\text{b}} (\text{C}_2\text{H}_5\text{COO}^{\text{-}}) \text{ is } 9.13 \text{ and } K_{\text{b}} = 10^{-9.13} = 7.41 \times 10^{-10}$$

$$7.41 \times 10^{-10} = [\text{OH}^{\text{-}}]^2 / 0.0175 \text{ so } [\text{OH}^{\text{-}}] = 3.60 \times 10^{-6} \text{ mol L}^{-1}$$

$$\text{pH} = 8.56$$


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- (c) What is meant by the term buffer solution?

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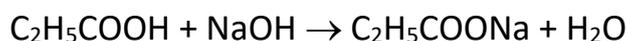
A solution which minimises changes in pH on addition of small amounts of acid or alkali

Significant quantities of weak acid and salt (a source of its conjugate base) are both present.

The weak acid combines with added  $\text{OH}^-$  ions. The conjugate base combines with added  $\text{H}_3\text{O}^+$  ions.

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- (d) When a solution of propanoic acid is partially neutralised using sodium hydroxide a buffer solution is made.



Calculate the pH of the buffer solution formed when 15.0 mL of a 0.250 mol L<sup>-1</sup> solution of sodium hydroxide is added to 25.0 mL of a 0.200 mol L<sup>-1</sup> solution of propanoic acid.

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Initial moles of  $\text{C}_2\text{H}_5\text{COOH} = CV = 0.200 \times 0.0250 = 5.00 \times 10^{-3}$  mol

Moles of NaOH added =  $CV = 0.250 \times 0.0150 = 3.75 \times 10^{-3}$  mol

Moles of  $\text{C}_2\text{H}_5\text{COONa}$  formed =  $3.75 \times 10^{-3}$  mol

Moles of  $\text{C}_2\text{H}_5\text{COOH}$  unreacted =  $5.00 \times 10^{-3} - 3.75 \times 10^{-3} = 1.25 \times 10^{-3}$

Total volume of buffer = 40.0 mL = 0.0400 L

Concentration of  $\text{C}_2\text{H}_5\text{COONa} = n/v = 3.75 \times 10^{-3} / 0.0400 = 0.0938$  mol L<sup>-1</sup>

Concentration of  $\text{C}_2\text{H}_5\text{COOH} = n/v = 1.25 \times 10^{-3} / 0.0400 = 0.0313$  mol L<sup>-1</sup>

$\text{pH} = \text{pK}_a + \log [\text{base}]/[\text{acid}]$

$\text{pH} = 4.87 + \log [0.0938]/[0.0313]$  pH = 5.35

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Note: Or can use  $[\text{H}_3\text{O}^+] = K_a \times ([\text{acid}]/[\text{salt}])$

Since we need ratio acid:base or acid:salt there is no need to convert the mol to concentration in either method

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**Question Three**

In a titration, 24.0 mL of sodium hydroxide solution was added, in 1 mL portions, to 20.0 mL of methanoic acid solution, HCOOH.

$$pK_a(\text{HCOOH}) = 3.75$$

After the addition of each 1 mL, the pH was measured and recorded using a pH meter.

- The concentration of the methanoic acid solution was  $0.400 \text{ mol L}^{-1}$ .
- The concentration of the sodium hydroxide solution was  $0.500 \text{ mol L}^{-1}$ .

(a) (i) Write the  $K_a$  expression for methanoic acid.

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$$K_a = \frac{[\text{HCOO}^-(\text{aq})] [\text{H}_3\text{O}^+(\text{aq})]}{[\text{HCOOH}(\text{aq})]}$$


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(ii) Calculate the pH of the  $0.400 \text{ mol L}^{-1}$  methanoic acid solution before the titration starts.

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$$K_a = 10^{-pK_a} = 1.78 \times 10^{-4} \quad K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HCOOH}]}$$

$$1.78 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+]^2}{0.400} \quad \text{so } [\text{H}_3\text{O}^+] = 8.43 \times 10^{-3} \text{ mol L}^{-1}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = 2.07$$


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(b) Show that the volume of sodium hydroxide solution required to reach the end point is 16.0 mL



$$n(\text{HCOOH}) = CV = 0.400 \times 0.0200 = 8.00 \times 10^{-3} \text{ mol}$$

This will react with  $8.00 \times 10^{-3} \text{ mol}$  of NaOH (as react in a 1:1 ratio)

$$V(\text{NaOH}) = n/C = 8.00 \times 10^{-3} / 0.500 = 0.0160 \text{ L} = 16.0 \text{ mL}$$


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- (c) By considering the amount of excess alkali remaining, calculate the pH of the solution after 24.0 mL of sodium hydroxide is added.

$$\text{Excess NaOH} = 24.0 - 16.0 = 8.00 \text{ mL.}$$

$$n(\text{NaOH}) = 0.500 \times 0.00800 = 4.00 \times 10^{-3} \text{ mol}$$

This is now in a total volume of 44.0 mL

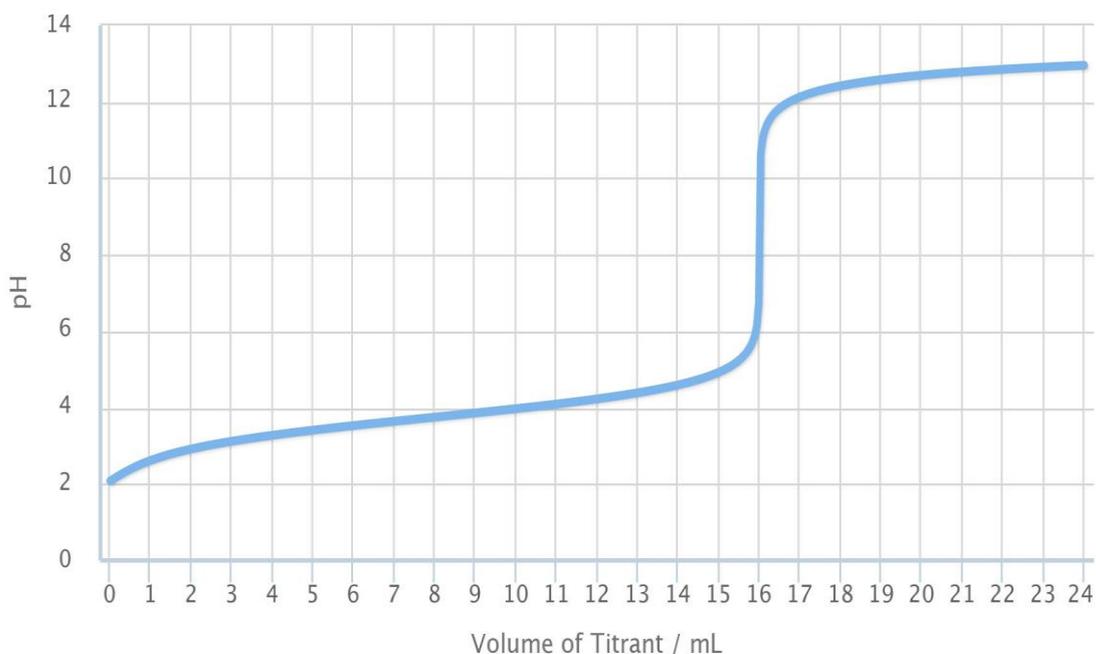
$$C(\text{NaOH}) = n/V = 4.00 \times 10^{-3} / 0.0440 = 0.0909 \text{ mol L}^{-1}$$

$$[\text{OH}^-] = 0.0909 \text{ therefore pH} = (12.96) \text{ } 13.0$$

- (d) (i) Sketch a graph to show how the pH changes during the titration.

S-shaped curve with kink at start. Start pH  $\approx$  2. end pH  $\approx$  13. pH at 8 mL (half way to EP) will equal pKa of 3.75. Vertical region at 16 mL. centred at pH between 8-9 (clearly at pH  $>$  7) and should start approx.. around 6 and end around 11.

pH Curve



- (ii) Using your sketch graph, justify why bromocresol green indicator ( $pK_a = 4.7$ ) *would or would not* be a suitable indicator for this titration.
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Bromocresol green indicator would be unsuitable as an indicator.

The pH range of an indicator =  $pK_a \pm 1$  With a  $pK_a$  of 4.7, the visible colour change would be in pH range of approx. 3.7-5.7 and would change colour at a volume well before the equivalence point.

To be suitable the range must fall in the vertical portion of the titration curve / where pH changes sharply so you would choose an indicator with a  $pK_a$  that is closest to the pH at the equivalence point of the titration.

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