

9 Four beakers each contain equal volumes of four solutions.

Beaker 1 contains 0.01 mol dm^{-3} NaOH

Beaker 2 contains 0.01 mol dm^{-3} NH_3

Beaker 3 contains 0.1 mol dm^{-3} NaOH

Beaker 4 contains 0.1 mol dm^{-3} $\text{Ba}(\text{OH})_2$

The pH of the four solutions was measured.

Which of the following gives the order of **decreasing** pH?

- A 1, 2, 3, 4
- B 2, 1, 3, 4
- C 3, 4, 1, 2
- D 4, 3, 1, 2

(Total for Question 9 = 1 mark)

8 At 100°C , pure water has a pH of 6, but at 25°C it has a pH of 7.

This is because

- A the dissociation of water is endothermic, so the concentration of hydrogen ions is lower at 100°C than at 25°C
- B the dissociation of water is exothermic, so the concentration of hydrogen ions is lower at 100°C than at 25°C
- C the dissociation of water is endothermic, so the concentration of hydrogen ions is higher at 100°C than at 25°C
- D the dissociation of water is exothermic, so the concentration of hydrogen ions is higher at 100°C than at 25°C

(Total for Question 8 = 1 mark)

7 At 50°C , the ionic product of water, K_w , is $5.5 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$.

At this temperature, water is

- A neutral with a pH of 7.0
- B neutral with a pH of 6.6
- C acidic with a pH of 6.6
- D alkaline with a pH of 7.4

(Total for Question 7 = 1 mark)

11 What is the pH of a $0.100 \text{ mol dm}^{-3}$ solution of sodium hydroxide at 283 K?

[$pK_w = 14.53$ at 283 K]

- A 13.00
- B 13.53
- C 14.43
- D 14.53

5: At 25°C , the pH of pure water is 7.0 and at 75°C the pH is 6.4 .

Why is the pH of pure water lower at 75°C than at 25°C ?

- A there are more hydrogen ions than hydroxide ions when the water is at 75°C
- B the dissociation of water is endothermic so the concentration of hydrogen ions is lower at 75°C
- C the dissociation of water is endothermic so the concentration of hydrogen ions is higher at 75°C
- D the dissociation of water is exothermic so the concentration of hydrogen ions is higher at 75°C

11 A 1 mol dm^{-3} solution of ethanoic acid is gradually diluted by the addition of distilled water.

What happens to the degree of dissociation of the acid and the pH of the solution as the distilled water is added?

	Degree of dissociation	pH
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(Total for Question 11 = 1 mark)

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19 This question is about acids and bases.

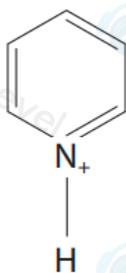
(a) Some K_a and pK_a values for several weak acids are shown.

Weak acid	$K_a / \text{mol dm}^{-3}$	pK_a
$\text{C}_5\text{H}_5\text{NH}^+$		5.25
$\text{CH}_3\text{CH}_2\text{COOH}$	1.3×10^{-5}	4.88
HCOOH	1.8×10^{-4}	3.75
CH_2ClCOOH	1.4×10^{-3}	2.85
CHCl_2COOH	4.5×10^{-2}	

(i) Complete the table.

(2)

(ii) The structure of $\text{C}_5\text{H}_5\text{NH}^+$ is shown.



Write the K_a expression for this weak acid.

(1)

(iii) When methanoic acid is added to propanoic acid, an equilibrium is set up containing two acid-base pairs.

Complete the equilibrium, labelling the acid-base pairs as **A1**, **B1** and **A2**, **B2**.

(2)



- *(iv) A student measured the pH of separate $0.050 \text{ mol dm}^{-3}$ solutions of CH_2ClCOOH and of CHCl_2COOH , using a pH meter.

The student also calculated the pH of these solutions, making two assumptions:

- $[\text{H}^+] = [\text{A}^-]$
- $[\text{HA}]_{\text{equilibrium}} = [\text{HA}]_{\text{initial}}$

The measured and calculated pH values are shown.

	$0.050 \text{ mol dm}^{-3} \text{ CH}_2\text{ClCOOH}$	$0.050 \text{ mol dm}^{-3} \text{ CHCl}_2\text{COOH}$
Measured pH value	2.11	1.52
Calculated pH value	2.08	1.32

Discuss the differences between the student's measured and calculated pH values.

(6)

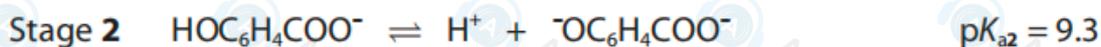
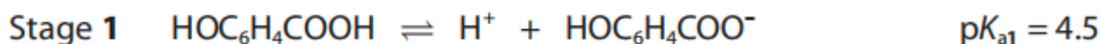
In your answer, you should

- show how the student calculated their pH values

$$[K_a(\text{CH}_2\text{ClCOOH}) = 1.4 \times 10^{-3} \text{ mol dm}^{-3}; \\ K_a(\text{CHCl}_2\text{COOH}) = 4.5 \times 10^{-2} \text{ mol dm}^{-3}]$$

- explain, with reference to the assumptions made, why there is a difference between the calculated and measured pH values
 - suggest why the measured pH values are higher than the calculated pH values.
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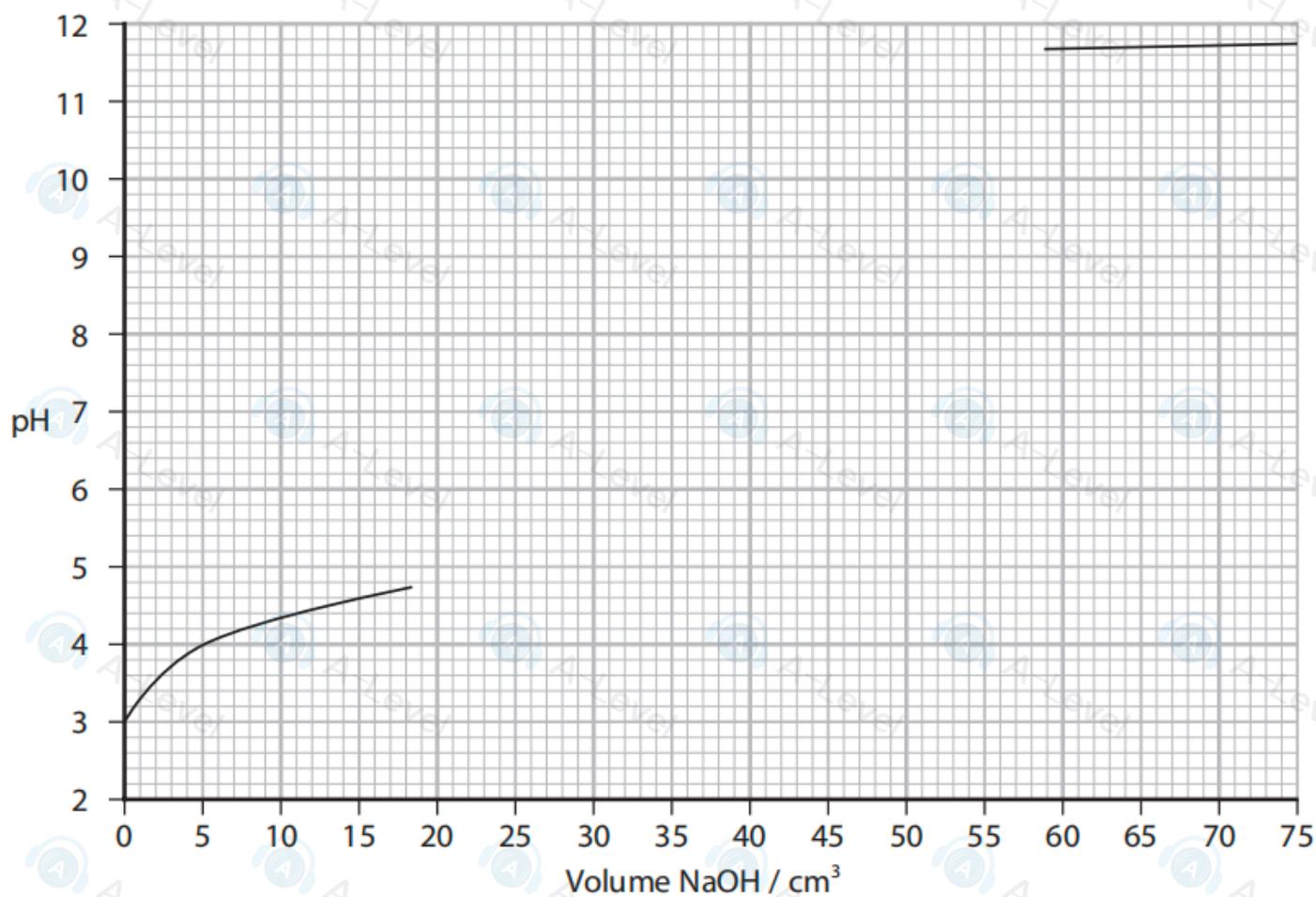
(b) 4-Hydroxybenzoic acid, $\text{HOC}_6\text{H}_4\text{COOH}$, is a diprotic acid, which dissociates in two stages. Each stage has a different $\text{p}K_{\text{a}}$ value.



In a titration, 75.0cm^3 of 0.025mol dm^{-3} NaOH was added to 25.0cm^3 of 0.025mol dm^{-3} $\text{HOC}_6\text{H}_4\text{COOH}$.

Complete the titration curve.

(3)



(c) A student was given 50.0 cm^3 of a solution of sodium hydroxide.

The pH of this solution was 12.43.

The student was asked to adjust the pH to 12.00, by dilution with deionised water. The student did **not** have access to a pH meter.

Calculate the volume of deionised water, **in cm^3** , the student should add to the original solution.

(5)

17: This question is about weak acids.

The values for the acid dissociation constant of ethanoic acid and chloroethanoic acid are shown.

Acid	$K_a / \text{mol dm}^{-3}$
ethanoic acid	1.70×10^{-5}
chloroethanoic acid	1.30×10^{-3}

(a) (i) Write the expression for K_a for ethanoic acid, CH_3COOH .

(1)

(ii) Calculate the pH of a 0.25 mol dm^{-3} solution of ethanoic acid.

(2)

(iii) State **one** assumption you have made when calculating the pH in (a)(ii).

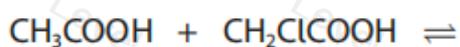
(1)

(b) (i) Suggest a reason why chloroethanoic acid, CH_2ClCOOH , is a stronger acid than ethanoic acid.

(1)

- (ii) When ethanoic acid and chloroethanoic acid are added together an equilibrium is set up containing two acid-base pairs.

Complete the equilibrium equation, labelling the conjugate acid-base pairs as A1, B1 and A2, B2.



(2)

- (c) A buffer was made by mixing 8.20 g of sodium ethanoate, CH_3COONa , with 250 cm^3 of $0.700 \text{ mol dm}^{-3}$ solution of ethanoic acid, CH_3COOH .

Calculate the pH of the buffer.

(4)

- (d) A type of bubble bath contains a weak acid and the indicator bromocresol green. The bubble bath is initially yellow but when diluted with bath water, its colour changes from yellow to green to blue.

Use the information on page 9 of the Data Booklet to deduce why the bubble bath changes colour.

(3)
