Acids and bases

pH and K_w

1. Complete the table below showing some numbers and their common logarithmic values; (1 mark)

Number (*n*)	**log$_{10}$ *n***
0.001	………..
0.1	………..
………..	0
………..	3

2. Calculate the pH (to 2 dp) of each of the solutions below: (3 marks)

(a) $[H^+] = 1.00 \times 10^{-10}$ mol dm^{-3} pH = ...

(b) $[H^+] = 0.200 \times 10^{-2}$ mol dm^{-3} pH = ...

(c) $[H^+] = 3.50 \times 10^{-3}$ mol dm^{-3} pH = ...

3. As water is always slightly ionised, we can write the following equilibrium for water;

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \quad \Delta H = +ive$$

As only a very small amount of the water is ionised, we define a new equilibrium constant for this equilibrium called the **ionic product of water**, K_w;

$$K_w = [H^+(aq)]\,[OH^-(aq)]$$

Like any other equilibrium constant, the value of K_w depends on the temperature of the equilibrium.

(a) Predict what effect increasing the temperature will have **on the pH** of pure water. (1 mark)

(b) Calculate the pH of pure water (to 2 dp) at each of the temperatures below;

(i) 10 °C, $K_w = 0.29 \times 10^{-14}$ mol^2 dm^{-6}

(ii) 25 °C, $K_w = 1.01 \times 10^{-14}$ mol^2 dm^{-6}

(iii) 40 °C, $K_w = 2.92 \times 10^{-14}$ mol^2 dm^{-6}

(3 marks)

(c) Complete the paragraph below;

As the temperature decreases, water becomes (more acidic / less acidic / remains neutral). Explain your answer

(2 marks)

pH and acids

1. Identify the species formed when the following act as acids;

(a) HCl ______________________________

(b) NH_4^+______________________________

(c) HCO_3^-______________________________

(3 marks)

2. Calculate the pH (to 2 dp) of the following acids;

(a) 0.25 mol dm^{-3} HCl ______________________________ (1 mark)

(b) 0.004 mol dm^{-3} $NaHSO_4$, K_a of HSO^-_4 = 1.00×10^{-2} mol dm^{-3}

(2 marks)

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3. Calculate the concentration of the following acids given their pH.

(a) HCl, pH 0.65 ______________________________ (1 mark)

(b) H_2SO_4, pH 2.61______________________________ (1 mark)

(c) CH_3COOH, pH 3.40, K_a 1.7×10^{-5} mol dm^{-3}

(2 marks)

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pH and bases

1. Define;

(a) a Brønsted-Lowry acid

(1 mark)

(b) a Brønsted-Lowry base

(1 mark)

2. In the following acid-base reactions identify the **reactant species** (ion or molecule) acting as a Brønsted-Lowry base;

(a) $2\ NH_3 + H_2SO_4 \rightarrow (NH_4^+)_2SO_4^{2-}$
(b) $Ca(OH)_2 + H_2CO_3 \rightarrow CaCO_3 + 2\ H_2O$
(c) $Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2O + CO_2$

(3 marks)

3. Calculate the pH (to 2 dp) of the following basic solutions (take K_w to be 1.00×10^{-14} mol^2 dm^{-6});

(a) 0.150 mol dm^{-3} NaOH

(1 mark)

(b) 0.261 mol dm^{-3} $Mg(OH)_2$

(1 mark)

4. Calculate the concentration of the following basic solutions (take K_w to be 1.00×10^{-14} mol^2 dm^{-6});

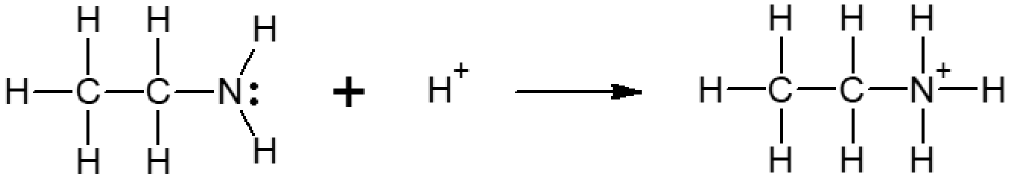
(a) KOH, pH 11.00

(1 mark)

(b) $Ca(OH)_2$, pH 10.45

(1 mark)

5. Ethylamine is a weak base. Draw a curly arrow on the diagram below to show how the ethylamine acts as a base.



Acid-base tritrations

Some students are carrying out an investigation into the neutralisation reactions between strong acids and bases and weak acids and bases.

They titrate 25 cm^3 samples of four different bases against four different acids as shown in the table below.

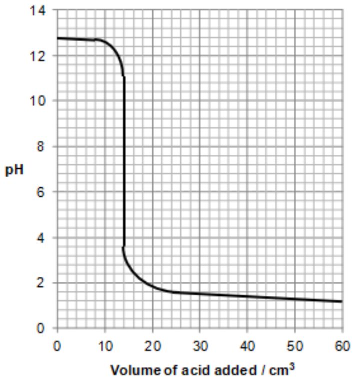
For each of the titrations **1 - 4**;

(a) Choose the correct titration curve from those shown below,

(b) Name a suitable indicator for the titration,

(c) For titrations **1** and **2**, calculate the concentration of the acid.

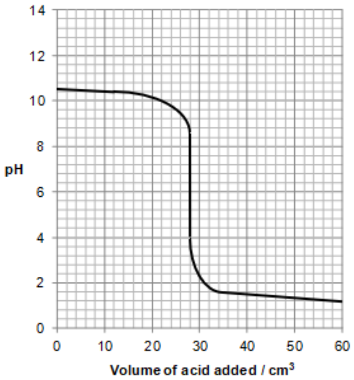
Titration	Base	Acid
1	0.100 mol dm^{-3} NaOH	HCl
2	0.100 mol dm^{-3} KOH	HCOOH
3	0.100 mol dm^{-3} NH_3 solution	HNO_3
4	0.100 mol dm^{-3} $NaHCO_3$	CH_3COOH



Titration number: 1 2 3 4

Suitable indicator: ....

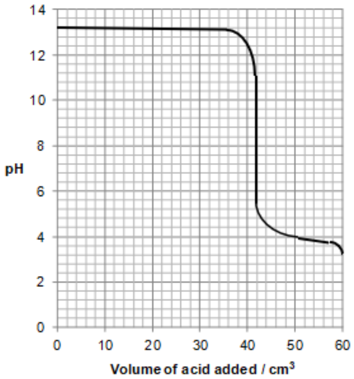
Conc. of acid (if needed): ....



Titration number: 1 2 3 4

Suitable indicator: ....

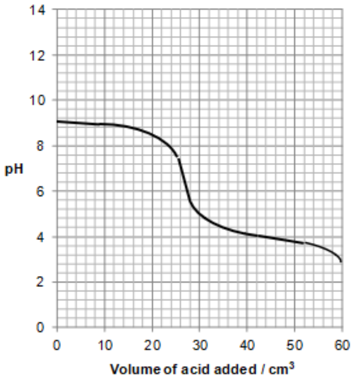
Conc. of acid (if needed): ....



Titration number: 1 2 3 4

Suitable indicator: ....

Conc. of acid (if needed):......................................



Titration number: 1 2 3 4

Suitable indicator: ....

Conc. of acid (if needed):......................................

Buffer solutions

A buffer solution is a solution that resists a change in pH when a small quantity of acid or base is added.

1. (a) A buffer solution is made by mixing 0.510 mol of methanoic acid with 0.450 mol of sodium methanoate in 500 cm^3 of water.

(i) Write an equation to represent the equilibrium established in the buffer solution.

(1 mark)

(ii) Calculate the pH of the buffer solution formed. (pK_a for methanoic acid = 3.75)

(3 marks)

(b) Explain how this buffer resists change in pH on;

(i) addition of a small quantity of acid.

(1 mark)

(ii) addition of a small quantity of base.

(1 mark)

2. Mark and Karen are carrying out a science project on the application of buffer solutions in the human body. They have discovered that a buffer of carbonic acid (H_2CO_3) and hydrogen carbonate (HCO_3^-) is present in blood plasma to maintain a pH of between 7.35 and 7.45.

(a) They would like to recreate a similar buffer solution in the laboratory. In what proportions should they mix 0.150 mol dm^{-3} solutions of carbonic acid and sodium hydrogen carbonate to give a buffer solution with a pH of 7.40? (K_a for H_2CO_3 is 4.5×10^{-7} mol dm^{-3}).

(2 marks)

(b) Why do you think buffer solutions are needed in the human body?

(2 marks)

More complex buffer calculations

Scientists wish to investigate whether certain bacteria can adapt to live in acidic conditions.

1. The scientists make up a buffer solution, by mixing 15.0 cm^3 of a

0.100 mol dm^{-3} aqueous solution of NaOH with 35.0 cm^3 of a

0.150 mol dm^{-3} solution of propanoic acid. Calculate the pH of the buffer solution formed.

(K_a for propanoic acid has the value 1.35×10^{-5} mol dm^{-3})

(4 marks)

2. The scientists wish to test if the solution formed is indeed a buffer solution and will resist change in pH on the addition of small quantities of acid or base possibly formed by the bacteria. They take two separate 10 cm^3 aliquots of the buffer solution formed in question **1** and add;

(a) 0.5 cm^3 of a 0.05 mol dm^{-3} solution of hydrochloric acid to one of the aliquots, and

(b) 0.5 cm^3 of a 0.05 mol dm^{-3} solution of calcium hydroxide to the other aliquot.

Calculate the pH of each of the new solutions formed.

(6 marks)

Acids and bases – Answers

1.

Number (*n*)	**$\log_{10} n$**
0.001	–3
0.1	–1
1	0
1,000	3

(1 mark for all numbers correct)

2. (a) pH = 10.00
(b) pH = 2.70
(c) pH = 2.46

(3 marks, 1 mark for each correct answer given to 2 dp)

3. (a) Ionisation of water is endothermic (ΔH +ive) so increasing the temperature will favour the forward reaction and hence the [H+(aq)] will increase. As a result the pH of the water will decrease as the temperature is increased.

(1 mark)

(b)
(i) 10 °C, $K_w = 0.29 \times 10^{-14}$ mol^2 dm^{-6}; $[H^+(aq)] = 5.39 \times 10^{-8}$ mol dm^{-3} ∴ pH = 7.27
(ii) 25 °C, $K_w = 1.01 \times 10^{-14}$ mol^2 dm^{-6}; $[H^+(aq)] = 1.00 \times 10^{-7}$ mol dm^{-3} ∴ pH = 7.00
(iii) 40 °C, $K_w = 2.92 \times 10^{-14}$ mol^2 dm^{-6}; $[H^+(aq)] = 1.71 \times 10^{-7}$ mol dm^{-3} ∴ pH = 6.77

(3 marks)

(c) As the temperature decreases, water remains neutral

(1 mark)

Water is always neutral as $[H^+(aq)] = [OH^-(aq)]$ and so there is always an equal number of H^+ ions and OH^- ions.

(1 mark)

pH and acids

1. (a) $HCl \rightarrow H^+ + \underline{Cl^-}$

(b) $NH_4^+ \rightleftharpoons H^+ + \underline{NH_3}$

(c) $HCO_3^- \rightleftharpoons H^+ + \underline{CO_3^{2-}}$

(3 marks)

2. (a) $pH = -\log[0.25] = 0.60$

(1 mark)

(b) $K_a = \dfrac{[H^+][SO_4^{2-}]}{[HSO_4^-]} = \dfrac{[H^+]^2}{[HSO_4^-]}$

$\therefore\ [H^+]^2 = (1.0 \times 10^{-2}) \times 0.004 = 4 \times 10^{-5}$

$[H^+] = 6.32 \times 10^{-3}\ \text{mol dm}^{-3}$

$\therefore\ pH = -\log[6.32 \times 10^{-3}]$

$\therefore\ \underline{pH\ \ = 2.20}$

(1 mark for K_a expression, 1 mark for pH)

3. (a) $[H^+] = 10^{-0.65}\ \therefore\ [H^+] = 0.22\ \text{mol dm}^{-3}\ \therefore\ \underline{[HCl] = 0.22\ \text{mol dm}^{-3}}$ *(1 mark)*

(b) $[H^+] = 10^{-2.61}\ \therefore\ [H^+] = 2.45 \times 10^{-3}\ \text{mol dm}^{-3}$

$H_2SO_4 \rightarrow 2\,H^+ + SO_4^{2-}$ and $\therefore\ [H_2SO_4] = [H^+]/2 = \underline{1.23 \times 10^{-3}\ \text{mol dm}^{-3}}$

(1 mark)

(c) $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-\ \therefore\ K_a = \dfrac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \dfrac{[H^+]^2}{[CH_3COOH]}$

$pH = -\log[H^+],\ \therefore\ [H^+] = 10^{-3.40},\ \therefore\ [H^+] = 3.98 \times 10^{-4}\ \text{mol dm}^{-3}$

(1 mark)

$\therefore\ 1.7 \times 10^{-5} = [3.98 \times 10^{-4}]^2 / [CH_3COOH]$

$\therefore\ \underline{[CH_3COOH] = 9.32 \times 10^{-3}\ \text{mol dm}^{-3}}$

(1 mark)

pH and bases

1. (a) A Brønsted-Lowry acid is a proton donor

(1 mark)

(b) A Brønsted-Lowry base is a proton acceptor

(1 mark)

2. (a) $2\ NH_3 + H_2SO_4 \rightarrow (NH_4^+)_2\ SO_4^{2-}$ Basic species = NH_3

(b) $Ca(OH)_2 + H_2CO_3 \rightarrow CaCO_3 + 2\ H_2O$ Basic species = OH^-

(c) $Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2O + CO_2$ Basic species = CO_3^{2-}

(3 marks)

3. (a) $[NaOH] = 0.150\ mol\ dm^{-3}$ and $\therefore\ [OH^-] = 0.150\ mol\ dm^{-3}$

$\therefore\ 1 \times 10^{-14} = [H^+][0.150]$ and so, $[H^+] = 6.67 \times 10^{-14}\ mol\ dm^{-3}$

$\therefore\ pH = -\log[6.67 \times 10^{-14}] = \underline{13.18}$

(1 mark)

(b) $[Mg(OH)_2] = 0.261\ mol\ dm^{-3}$ and $\therefore\ [OH^-] = 0.261 \times 2 = 0.522\ mol\ dm^{-3}$

$\therefore\ 1 \times 10^{-14} = [H^+][0.522]$ and so, $[H^+] = 1.92 \times 10^{-14}\ mol\ dm^{-3}$

$\therefore\ pH = -\log[1.92 \times 10^{-14}] = \underline{13.72}$

(1 mark)

4. (a) $11.00 = -\log[H^+]$ and $\therefore\ [H^+] = 1.00 \times 10^{-11}\ mol\ dm^{-3}$

$\therefore\ 1.00 \times 10^{-14} = [1.00 \times 10^{-11}][OH^-]$ and so, $[OH^-] = 1.00 \times 10^{-3}\ mol\ dm^{-3}$

$\therefore\ \underline{[KOH] = 1.00 \times 10^{-3}\ mol\ dm^{-3}}$

(1 mark)

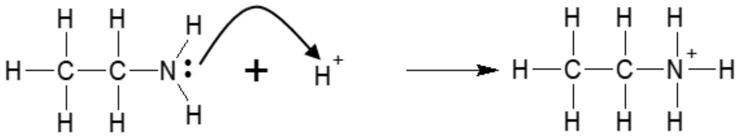
(b) $10.45 = -\log[H^+]$ and $\therefore\ [H^+] = 3.55 \times 10^{-11}\ mol\ dm^{-3}$

$\therefore\ 1.00 \times 10^{-14} = [3.55 \times 10^{-11}][OH^-]$ and so, $[OH^-] = 2.82 \times 10^{-4}\ mol\ dm^{-3}$

Since 1 mol of $Ca(OH)_2$ contains 2 mol OH^-, $[Ca(OH)_2] = 1.41 \times 10^{-4}\ mol\ dm^{-3}$

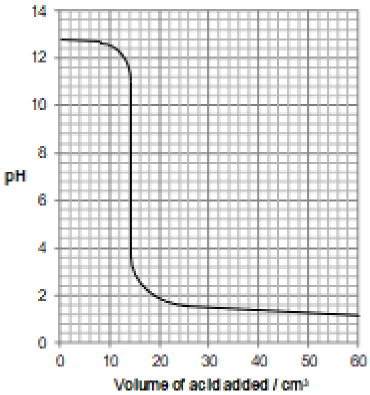
(1 mark)

5.



(1 mark)

Acid-base titrations



Titration number: [✓ 1] 2 3 4

Suitable indicator: phenolphthalein
or methyl orange

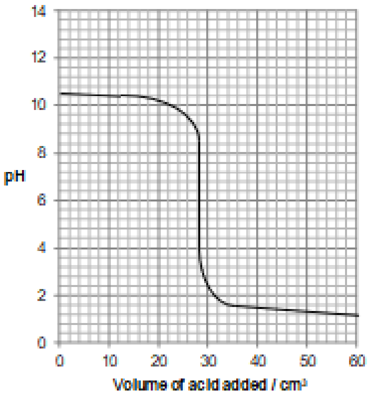
Conc. of acid (if needed):

$NaOH + HCl \rightarrow NaCl + H_2O$

Moles in 25 cm^3 0.100 mol dm^{-3} NaOH = 2.5×10^{-3}

Volume of HCl needed for neutralisation = 14 cm^3

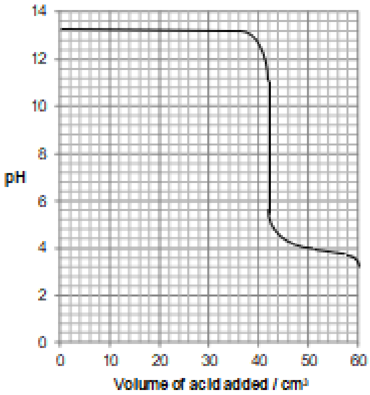
Conc. of HCl = $2.5 \times 10^{-3} / 0.014\ dm^3 = \underline{0.18\ mol\ dm^{-3}}$



Titration number: 1 2 [✓ 3] 4

Suitable indicator: methyl orange

Conc. of acid (if needed): N/A



Titration number: 1 [✓ 2] 3 4

Suitable indicator: phenolphthalein

Conc. of acid (if needed):

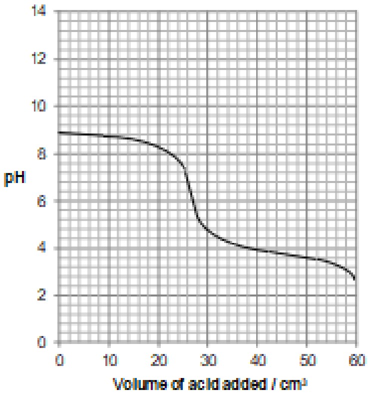
$KOH + HCOOH \rightarrow HCOO^-K^+ + H_2O$

Moles in 25 cm^3 0.100 mol dm^{-3} KOH = 2.5×10^{-3}

Volume of HCOOH needed for neutralisation = 42 cm^3

Conc. of HCOOH = $2.5 \times 10^{-3} / 0.042\ dm^3$

$= \underline{0.060\ mol\ dm^{-3}}$



Titration number: 1 2 3 [✓ 4]

Suitable indicator: none

Conc. of acid (if needed): N/A

(1 mark for correct identification of each titration,
1 mark for each suitable indicator named,
1 mark for each calculation of acid concentration)

Buffer solutions

1. (a) (i) $HCOOH(aq) \rightleftharpoons HCOO^-(aq) + H^+(aq)$

(1 mark)

(ii) $pK_a = -\log K_a$, $\therefore$ $\underline{K_a = 10^{-3.75} = 1.78 \times 10^{-4}\ mol\ dm^{-3}}$

(1 mark)

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

$[HCOO^-(aq)] = 0.450\ mol / 0.5\ dm^3 = 0.90\ mol\ dm^{-3}$
$[HCOOH(aq)] = 0.510\ mol / 0.5\ dm^3 = 1.02\ mol\ dm^{-3}$
Substituting these values in we get, $1.78 \times 10^{-4}\ mol\ dm^{-3} = 0.90 \times [H^+(aq)] / 1.02$
$\therefore$ $[H^+(aq)] = \underline{2.02 \times 10^{-4}\ mol\ dm^{-3}}$

(1 mark)

$\therefore$ $\underline{pH = 3.70}$

(1 mark)

(b) (i) On the addition of H^+ ions, according to Le Châtelier's principle, the equilibrium shifts to the left to remove the extra H^+ ions added and maintain the pH approximately constant.

(1 mark)

(ii) On the addition of OH^- ions, the OH^- ions react with the HCOOH to produce water molecules and more $HCOO^-$;

$$HCOOH + OH^- \rightarrow HCOO^- + H_2O$$

This removes the OH^- and so the pH remains approximately constant.

(1 mark)

2. (a) $H_2CO_3(aq) \rightleftharpoons HCO^-_3(aq) + H^+(aq)$
pH of desired buffer = 7.40, so $[H^+(aq)] = 10^{-7.40} = 3.98 \times 10^{-8}\ mol\ dm^{-3}$

(1 mark)

$$K_a = \frac{[HCO_3^-(aq)][H^+(aq)]}{[H_2CO_3(aq)]}$$

$$\therefore \quad \frac{[HCO_3^-(aq)]}{[H2CO3(aq)} = \frac{K}{[H+(aq)]} = \frac{4.5 \times 10^{-7}\ mol\ dm^{-3}}{3.98 \times 10{-}8\ mol\ dm{-}3} = \frac{11.3}{1}$$

Since both stock solutions are of an equal concentration they should mix the two in a ratio of $\underline{11.3 : 1\ HCO_2^- : H_2CO_3}$

(b) Many reactions in the human body rely on enzymes. Enzymes work only under very precise conditions. If the pH moves outside of a narrow range, the enzymes slow or stop working and can be denatured. Hence maintaining a constant pH is essential.

(2 marks)

More complex buffer calculations

1. $CH_3CH_2COOH + NaOH \rightarrow CH_3CH_2COO^-Na^+ + H_2O$

Moles of NaOH = $0.015\ dm^3 \times 0.100\ mol\ dm^{-3} = 1.5 \times 10^{-3}$ mol

(1 mark)

∴ moles of CH_3CH_2COOH will decrease by 1.5×10^{-3} mol and moles of $CH_3CH_2COO^-Na^+$ will increase by 1.5×10^{-3} mol.

(1 mark)

	CH_3CH_2COOH	$\rightleftharpoons$	*$CH_3CH_2COO^-$*	*+ H^+*
***I**nitial moles*	*$0.035\ dm^3 \times 0.150\ mol\ dm^{-3}$* *$= 5.25 \times 10^{-3}$ mol*		*0 mol*	*0 mol*
***C**hange in moles*	*-1.5×10^{-3} mol*		*$+1.5 \times 10^{-3}$ mol*	*?*
***E**quilibrium moles*	*3.75×10^{-3} mol*		*1.5×10^{-3} mol*	*?*

$$K_a = \frac{[CH_3CH_2COO^-][H^+]}{[CH_3CH_2COOH]} = \frac{(1.5 \times 10^{-3}\ mol / 0.05\ dm^3) \times [H^+]}{(3.75 \times 10^{-3}\ mol / 0.05\ dm^3)} = 1.35 \times 10^{-5}\ mol\ dm^{-3}$$

∴ $[H^+] = 3.38 \times 10^{-5}\ mol\ dm^{-3}$

∴ <u>pH = 4.47</u>

(1 mark)

2. In a 10 cm³ aliquot (= 1/5 th) of the buffer solution made above;

moles of $CH_3CH_2COOH = 7.5 \times 10^{-4}$ mol; moles of $CH_3CH_2COO^- = 3.0 \times 10^{-4}$ mol

(a) No. of moles of acid added = $0.0005\ dm^3 \times 0.05\ mol\ dm^{-3} = 2.5 \times 10^{-5}$ mol

$CH_3CH_2COO^- + H^+ \rightarrow CH_3CH_2COOH$

∴ moles of CH_3CH_2COOH will increase by 2.5×10^{-5} mol and moles of $CH_3CH_2COO^-$ will decrease by 2.5×10^{-5} mol.

(1 mark)

	CH_3CH_2COOH	$\rightleftharpoons$	$CH_3CH_2COO^-$	+ H^+
Initial moles	7.5×10^{-4} mol		3.0×10^{-4} mol	
Change in moles	$+2.5 \times 10^{-5}$ mol		-2.5×10^{-5} mol	?
Equilibrium moles	7.75×10^{-4} mol		2.75×10^{-4} mol	?

$$\therefore\ 1.35 \times 10^{-5}\ mol\ dm^{-3} = \frac{(2.75 \times 10^{-4}\ mol / 0.0105\ dm^3) \times [H^+]}{(7.75 \times 10^{-4}\ mol / 0.0105\ dm^3)}$$

∴ $[H^+] = 3.80 \times 10^{-5}\ mol\ dm^{-3}$

∴ <u>pH = 4.42</u>

(1 mark)

(b) No. of moles of $Ca(OH)_2$ added = $0.0005\ dm^3 \times 0.05\ mol\ dm^{-3} = 2.5 \times 10^{-5}$ mol

∴ no. of moles of OH^- added = $2 \times 2.5 \times 10^{-5}$ mol = 5.0×10^{-5} mol

(1 mark)

$$CH_3CH_2COOH + OH^- \rightarrow CH_3CH_2COO^- + H_2O$$

∴ moles of CH_3CH_2COOH will decrease by 5.0×10^{-5} mol and moles of $CH_3CH_2COO^-$ will increase by 5.0×10^{-5} mol.

(1 mark)

	CH_3CH_2COOH	⇌	$CH_3CH_2COO^-$	+	H^+
Initial moles	7.5×10^{-4} mol		3.0×10^{-4} mol		
Change in moles	-5.0×10^{-5} mol		$+5.0 \times 10^{-5}$ mol		?
Equilibrium moles	7.0×10^{-4} mol		3.5×10^{-4} mol		?

$$\therefore 1.35 \times 10^{-5}\ mol\ dm^{-3} = \frac{(3.5 \times 10^{-4}\ mol / 0.0105\ dm^3) \times [H^+]}{(7.0 \times 10^{-4}\ mol / 0.0105\ dm^3)}$$

∴ $[H^+] = 2.7 \times 10^{-5}\ mol\ dm^{-3}$

∴ <u>pH = 4.57</u>

(1 mark)