At 40 °C the ionic product of water, $K_w = 2.92 \times 10^{-14}$ mol^2 dm^{-6}

(a) Give the expression for K_w

Calculate the pH of pure water at 40 °C
Give your answer to 2 decimal places.

K_w

pH ____________________

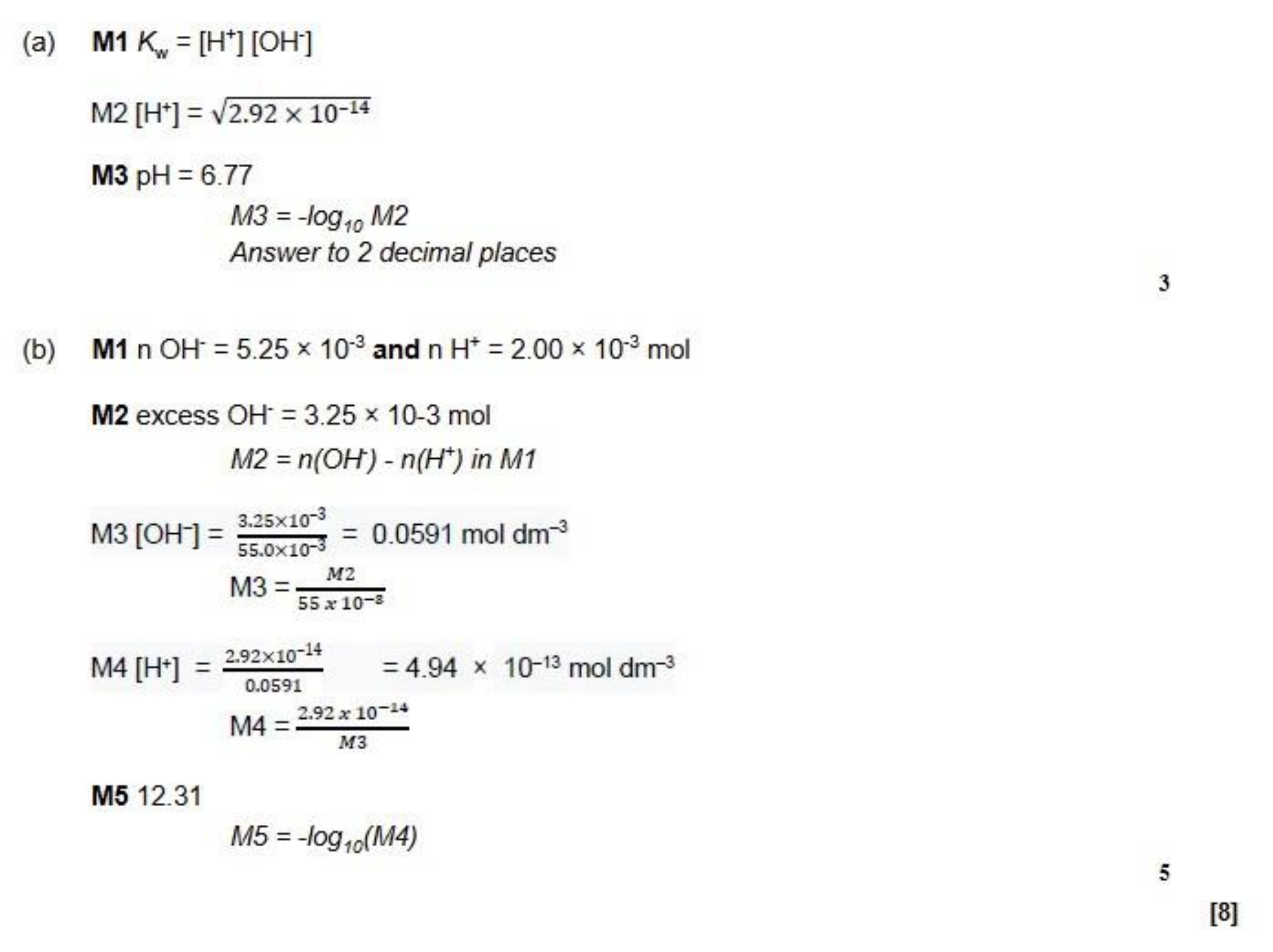
(3)

(b) 35.0 cm^3 of 0.150 mol dm^{-3} aqueous sodium hydroxide are mixed with 20.0 cm^3 of a 0.100 mol dm^{-3} solution of hydrochloric acid.
The temperature of the solution formed is 40 °C

Calculate the pH of the solution formed.
Give your answer to 2 decimal places.

pH ____________________

(5)



(a) **M1** $K_w = [H^+][OH^-]$

M2 $[H^+] = \sqrt{2.92 \times 10^{-14}}$

M3 pH = 6.77

M3 = -log₁₀ M2

Answer to 2 decimal places

3

(b) **M1** n OH^- = 5.25×10^{-3} **and** n H^+ = 2.00×10^{-3} mol

M2 excess OH^- = 3.25 × 10-3 mol

M2 = n(OH^-) - n(H^+) in M1

M3 $[OH^-] = \frac{3.25\times10^{-3}}{55.0\times10^{-3}} = 0.0591$ mol dm^{-3}

M3 = $\frac{M2}{55 \, x \, 10^{-8}}$

M4 $[H^+] = \frac{2.92\times10^{-14}}{0.0591} = 4.94 \times 10^{-13}$ mol dm^{-3}

M4 = $\frac{2.92 \, x \, 10^{-14}}{M3}$

M5 12.31

M5 = $-\log_{10}(M4)$

5

[8]

Question - 02

This question is about pH.

Pure water dissociates slightly.

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \qquad \Delta H = +57\ \text{kJ mol}^{-1}$$

The equilibrium constant, $K_c = \frac{[H^+][OH^-]}{[H_2O]}$

The ionic product of water, $K_w = [H^+][OH^-]$

(a) Explain why $[H_2O]$ is not shown in the K_w expression.

____

____

(1)

Table 1 shows how K_w varies with temperature.

Table 1

Temperature / °C	**K_w / mol^2 dm^{-6}**
10	2.93×10^{-15}
20	6.81×10^{-15}
25	1.00×10^{-14}
30	1.47×10^{-14}
50	5.48×10^{-14}

(b) Explain why the value of K_w increases as the temperature increases.

____

____

____

____

(2)

(c) Give the expression for pH.

Calculate the pH of pure water at 50 °C
Give your answer to 2 decimal places.

Explain why water is neutral at 50 °C

Expression ____

Calculation

pH ______________

Explanation ____

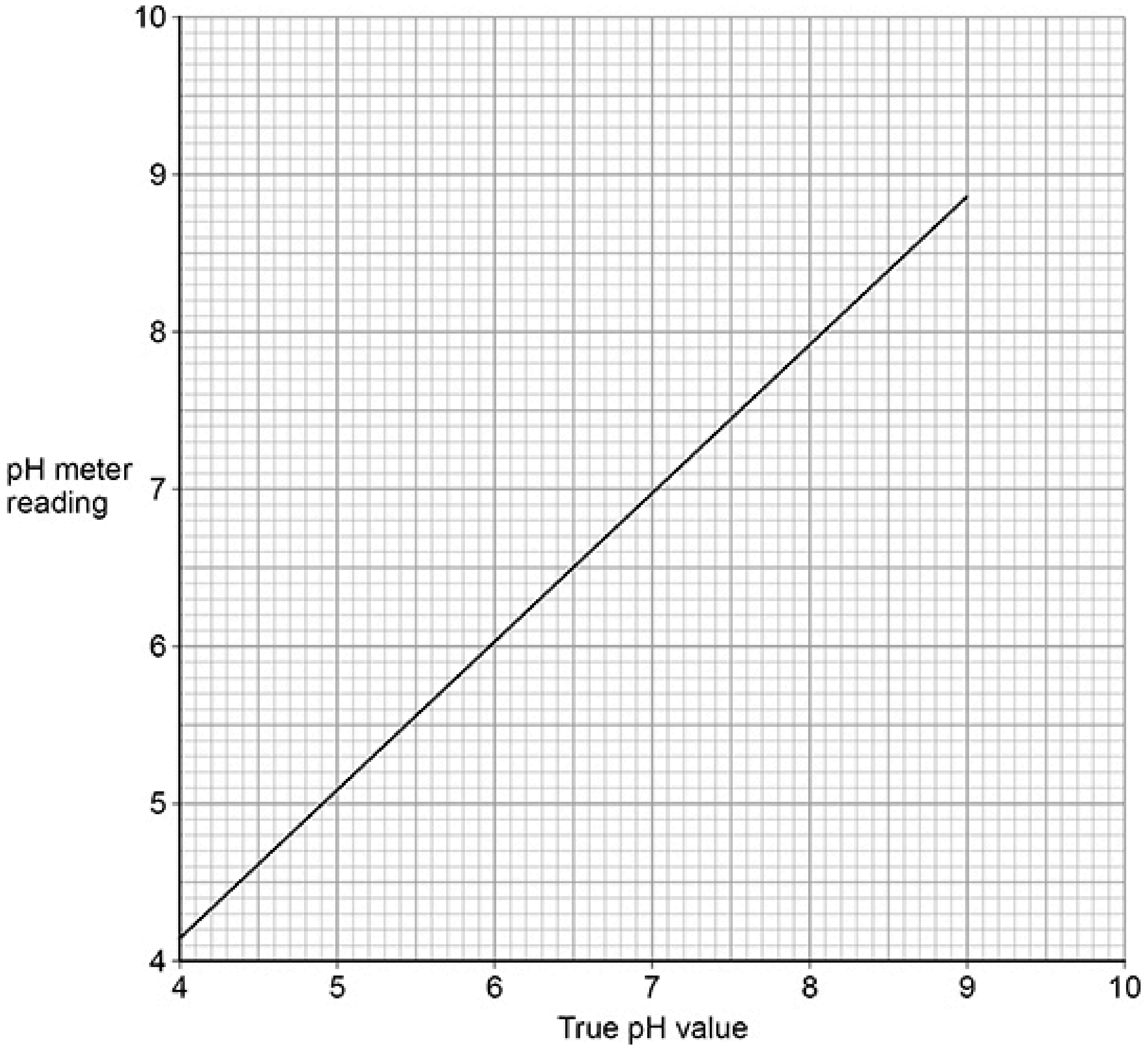
____

(4)

A pH meter is calibrated using a calibration graph.
To create the calibration, the pH meter is used to measure the pH of separate solutions, each with a known, accurate pH.

Figure 1 shows the calibration graph.

Figure 1



(d) Use **Figure 1** to give the true pH value when the pH meter reading is 5.6

____

(e) Suggest why the pH probe is washed with distilled water between each of the calibration measurements.

(1)

(f) The calibrated pH meter is used to monitor the pH during a titration of hydrochloric acid with sodium hydroxide.

Explain why the volume of sodium hydroxide solution added between each pH measurement is smaller as the end point of the titration is approached.

(1)

Figure 2 shows the pH curve for a titration of hydrochloric acid with sodium hydroxide solution.

Figure 2

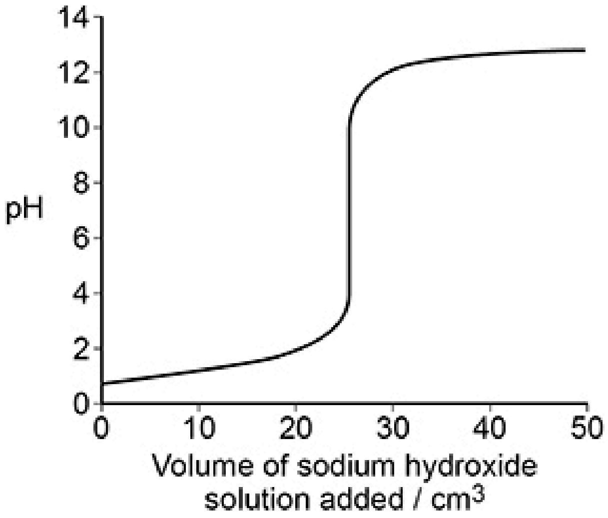


Table 2 shows data about some indicators.

Table 2

Indicator	**pH range**	**Colour at low pH**	**Colour at high pH**
Bromocresol green	3.8 – 5.4	yellow	blue
Phenol red	6.8 – 8.4	yellow	red
Thymolphthalein	9.3 – 10.5	colourless	blue

The student plans to do the titration again using one of the indicators in **Table 2** to determine the end point.

(g) State why all three of the indicators in **Table 2** are suitable for this titration.

(a) $[H_2O]$ is (almost) constant

Allow

$[H_2O]$ is (very) large in comparison (to $[H^+]$ and $[OH^-]$)

***or** $[H_2O]$ is incorporated in K_w*

***or** $K_w = K_c[H_2O]$*

***or** the equilibrium lies <u>very</u> much to the left.*

***Ignore** water has negligible dissociation*

***Ignore** $[H_2O] = 1$ or $[H_2O]$ is very small 1*

1

(b) M1 Equilibrium is endothermic (in forward direction)

1

M2 <u>Equilibrium</u> shifts to the RHS to minimise/oppose temperature increase

***Ignore** more H^+ and OH^- formed*

1

(c) M1 $pH = -\log_{10}[H^+]$

M1 Allow $pH = -\log[H^+]$

1

M2 $[H^+] = \sqrt{5.48 \times 10^{-14}}$ $(= 2.34 \times 10^{-7})$

M2 $[H^+]^2 = 5.48 \times 10^{-14}$

1

M3 $pH = -\log_{10} 2.34 \times 10^{-7} = 6.63$

M3 $pH = -\log_{10}$ M2

1

M4 $[H^+] = [OH^-]$

or

Dissociation of each water molecule gives one H^+ and one OH^-

M4 Allow equal amounts of H^+ and OH^-

1

(d) 5.55

Allow 5.5 to 5.6

1

(e) Different solutions must not contaminate each other

pH of previous solution doesn't contaminate new solution

or

To wash off any residual solution/substance (which could interfere with the reading)

Ignore to make neutral/neutralise

Ignore so as not to affect concentrations

1

(f) To avoid missing the end point

Or

(Very little pH change per cm^3 added at start) large change in pH (near end point)

1

(g) All have a colour change/pH range within the steep/vertical part of the titration curve

Colour change/pH range between pH 3 and 11

1

(h) M1 Amount of OH^- = $36.25 \times 0.200 \div 1000 = 7.25 \times 10^{-3}$ mol **and** Amount of H^+ = $25.0 \times 0.150 \div 1000 = 3.75 \times 10^{-3}$ mol

1

M2 Amount of excess OH^- = $7.25 \times 10^{-3} - 3.75 \times 10^{-3} = 3.50 \times 10^{-3}$ mol

1

M3 $[OH^-] = (3.50 \times 10^{-3}) \div (61.25 \times 10^{-3})$ (= 5.71×10^{-2} mol)

M3 $[OH^-] = (M2) \div (61.25 \times 10^{-3})$

1

M4 $[H^+] = \underline{1.00 \times 10^{-14}} \div 5.71 \times 10^{-2} = 1.75 \times 10^{-13}$

M4 $[H^+] = 1.00 \times 10^{-14} \div M3$

1

M5 pH = 12.76

M5 Allow pH = 12.8

M5 pH = $-\log_{10}(M4)$

Alternative Method

M4 p OH = 1.24

M5 pH = 14 – 1.24 = 12.76

1

End of Question

Question - 03

Equal volumes of two solutions, each with the same concentration, are mixed together at 298 K

Which two solutions, when mixed, form a solution with a pH >7?

A HCOOH and HCOOK ○

B KOH and CH_3COOH ○

C NH_3 and HCl ○

D NH_4Cl and KCl ○

B

KOH and CH_3COOH

Question - 04

A strong acid H_2X dissociates in aqueous solution.

$$H_2X(aq) \rightarrow 2\,H^+(aq) + X^{2-}\,(aq)$$

What is the pH of a 0.020 mol dm^{-3} solution of this acid?

A 1.00 ○

B 1.40 ○

C 1.70 ○

D 2.00 ○

B

1.40

Question - 05

Which can **not** function as a Brønsted-Lowry acid?

A CH_3COO^- ○

B HCO_3^- ○

C H_3O^+ ○

D NH_4^+ ○

A

CH_3COO^-

Question - 06

This question is about acids and bases.

(a) Calculate the pH of a 0.150 mol dm^{-3} solution of ethanoic acid at 25 °C
Give your answer to 2 decimal places.

For ethanoic acid, $K_a = 1.74 \times 10^{-5}$ mol dm^{-3} at 25 °C

pH ________________

(3)

(b) Strontium is an element in Group 2.

Calculate the pH of a 0.0100 mol dm^{-3} solution of strontium hydroxide at 10 °C
You may assume that strontium hydroxide is completely dissociated in this solution.

At 10 °C the ionic product of water, $K_w = 2.93 \times 10^{-15}$ mol^2 dm^{-6}

pH ________________

(3)

(c) The pH of a barium hydroxide solution is lower at 50 °C than at 10 °C

At 50 °C a 25 cm^3 sample of this barium hydroxide solution was neutralised by 22.45 cm^3 of hydrochloric acid added from a burette.

Deduce the volume of this hydrochloric acid that should be added from a burette to neutralise another 25 cm^3 sample of this barium hydroxide solution at 10 °C

Circle (◯) the correct answer.

> 22.45 cm^3 $= 22.45$ cm^3 < 22.45 cm^3

Explain your answer ____

____

____

____

(2)

(d) State how a buffer solution can be made from solutions of potassium hydroxide and ethanoic acid.

Give an equation for the reaction between potassium hydroxide and ethanoic acid.

State how this buffer solution resists changes in pH when a small amount of acid is added.

How buffer solution is made ____

____

____

____

Equation

____

How buffer solution resists pH change ___________________________________

____

____

____

(3)

(e) A buffer solution is made by adding 2.00 g of sodium hydroxide to 500 cm^3 of 1.00 mol dm^{-3} ethanoic acid solution.

Calculate the pH of this buffer solution at 25 °C

Give your answer to 2 decimal places.

For ethanoic acid, $K_a = 1.74 \times 10^{-5}$ mol dm^{-3} at 25 °C

pH ________________

(5)

(a) **M1** $K_a \dfrac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \dfrac{[H^+]^2}{[CH_3COOH]}$

1

M2 $[H^+] = \left(= \sqrt{1.74 \times 10^{-5} \times 0.150}\right) = 1.62 \times 10^{-3}$ (mol dm^{-3})

1

M3 pH = 2.79

M3 = −log M2

answer must be to 2 decimal places

1

(b) **M1** $[OH^-] = \underline{\mathbf{2 \times}}\ 0.01 = 0.02$

1

M2 $[H^+ \left(= \dfrac{K_w}{[OH^-]} = \dfrac{2.93 \times 10^{-15}}{0.02} \right) = 1.47 \times 10^{-13}$

allow $\dfrac{2.93 \times 10^{-15}}{M1}$

1

M3 pH = 12.83

allow 12.8

M3 = −log M2

1

if pH = 12.5(3) allow ***2*** *marks (not used factor of 2 in M1)*

(c) 22.45 or same

1

same $[OH^-]$ **or** amount/number of OH^- ions

1

(d) add excess ethanoic acid to KOH
or
add enough KOH to the ethanoic acid so that the acid is partially neutralised
or
add enough KOH so that the acid contains a mixture of ethanoic acid and ethanoate ions.

1

$KOH + CH_3COOH \longrightarrow CH_3COOK + H_2O$

allow $KOH + CH_3COOH \longrightarrow CH_3COO^-K^+ + H_2O$

1

CH_3COO^- (from salt) reacts with (added) acid/H^+

ignore equilibrium shifts

(e) **M1** (at start) n(NaOH) = $\frac{2.00}{40}$ = $\underline{0.05}$ (mol)

and

$n(CH_3COOH) = \frac{500 \times 1.0}{1000}$ = $\underline{0.5}$ (mol)

1

M2 (after adding NaOH)

$n(CH_3COOH)$ = (0.50 − 0.05) = 0.45 (mol)

1

M3 $n(CH_3COO^-)$ = (n NaOH) = 0.05 (mol)

M3 = n(NaOH) from M1

1

M4 $[H^+] = \left(\frac{K_a \times [CH_3COOH]}{[CH_3COO^-]} \text{ or } \frac{1.74 \times 10^{-5} \times 0.9}{0.1} \text{ or } \frac{1.74 \times 10^{-5} \times 0.45/V}{0.05/V} \right)$

= 1.57×10^{-4} (mol dm^{-3})

M4 = $\frac{1.74 \times 10^{-5} \times M2}{M3}$

V cancels out so not needed in expression

1

M5 pH = 3.80

answer to 2 decimal places

M5 allow 3.81

allow pH = −log M4

1

Henderson–Hasselbach method

M1 (at start) n(NaOH) = $\frac{2.00}{40}$ = $\underline{0.05}$ (mol)

and

$n(CH_3COOH) = \frac{500 \times 1.0}{1000}$ = $\underline{0.5}$ (mol)

1

M2 (after adding NaOH)

$n(CH_3COOH)$ = (0.50 − 0.05) = 0.45 (mol)

or $[CH_3COOH]$ = 0.9(0)

1

M3 $n(CH_3COO^-) = n(NaOH)$ = 0.05 (mol)

or $[CH_3COO^-$ = 0.1(0)

M3 = n(NaOH) from M1

1

M4 pH = 4.759 + log$\frac{[CH_3COO^-]}{[CH_3COOH]}$

or 4.759 + log$\frac{0.1}{0.9}$ **or** 4.759 + log$\frac{0.05/V}{0.45/V}$

V cancels out so not needed in this expression

1

Question - 07

Equal volumes of pairs of solutions are mixed.

Which pair forms a buffer solution?

A ammonia and ammonium chloride ○

B ammonia and methylamine ○

C ethanoic acid and methanoic acid ○

D hydrochloric acid and sodium hydroxide ○

A

ammonia and ammonium chloride

Question - 08

A 0.100 mol dm^{-3} solution of a weak acid has pH = 2.50

What is the value of K_a for this acid, in mol dm^{-3}?

A 3.16×10^{-2} ○

B 3.16×10^{-3} ○

C 1.00×10^{-4} ○

D 1.00×10^{-5} ○

C

1.00×10^{-4}

Question - 09

Which change causes the pH of 10 cm^3 of 1.0 mol dm^{-3} NaOH to be halved at 298 K?

$K_w = 1.0 \times 10^{-14}$ at 298 K

A adding 10 cm^3 of water ○

B adding 10 dm^3 of water ○

C adding 5 cm^3 of 1.0 mol dm^{-3} HCl ○

D adding 10 cm^3 of 1.0 mol dm^{-3} HCl ○

D

adding 10 cm^3 of 1.0 mol dm^{-3} HCl

Question - 10

Which species can behave as a Brønsted–Lowry acid in aqueous solution?

A SO_4^{2-} ○

B HCO_3^- ○

C BF_3 ○

D NH_3 ○

B

HCO_3^-

Calcium hydroxide is almost insoluble in water, but it reacts with dilute hydrochloric acid.

$$Ca(OH)_2(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + 2\ H_2O(l)$$

A student adds 100 cm^3 of 0.100 mol dm^{-3} hydrochloric acid to
0.600 g of solid calcium hydroxide.

(a) Show, by calculation, that the calcium hydroxide is in excess.

(2)

(b) The final mixture contains a saturated solution of $Ca(OH)_2$ at 293 K

At 293 K

- the solubility of $Ca(OH)_2$ in this solution is 0.400 g dm^{-3}
- $K_w = 6.80 \times 10^{-15}$ $mol^2\ dm^{-6}$

Calculate the pH of this solution.
Give your answer to two decimal places.

pH ____________________

(5)

(a) **M1** amount of HCl = 0.010 mol **AND**

amount $Ca(OH)_2$ (= $\frac{0.6}{74.1}$) = 0.00810

M2 0.00810 mol of $Ca(OH)_2$ requires 0.0162 mol HCl **OR**

0.01 mol of HCl requires 0.005 mol of $Ca(OH)_2$

NB There must be an indication that 0.005 is the amount of $Ca(OH)_2$ needed and not just 0.01/2

Alternative:

***M1** amount of HCl = 0.010 mol*

amount of $Ca(OH)_2$ needed = 0.0050 mol

***NB** There must be an indication that 0.005 is the amount of $Ca(OH)_2$ needed and not just 0.01/2*

***M2** mass of $Ca(OH)_2$ needed = 0.0050 x 74.1 = 0.3705 g (so 0.6 g is excess)*

2

(b) **M1** 0.400 g dm^{-3} means 0.400 ÷ 74.1 (= 0.00540 mol dm^{-3})

M2 $[OH^-]$ = **M1** x 2 (= 0.0108 mol dm^{-3})

M3 $[H^+] = 6.80 \times 10^{-15}$ ÷ **M2** (= 6.30×10^{-13} mol dm^{-3})

M4 pH = $-\log[H^+]$

M5 = -log **M3 with answer to 2dp** (= 12.20)

Correct answer to 2 dp = 5 marks

***ALLOW** 12.21 for 5 marks*

***M4 ALLOW** if calculation shown containing a number that has been calculated as $[H^+]$*

5

Question - 12

This question is about weak acids.

(a) The table below shows the pH ranges of some indicators.

Indicator	**pH range**
Bromocresol green	3.8 – 5.4
Bromothymol blue	6.0 – 7.6
Thymol blue	8.0 – 9.6

Identify the indicator that is most suitable for use in a titration between propanoic acid and sodium hydroxide.

____

(1)

(b) Give the expression for the acid dissociation constant (*K*a) for propanoic acid ($CH3CH_2COOH$).

*K*a

(1)

(c) Calculate the pH of a 0.100 mol dm^{-3} propanoic acid solution.
Give your answer to 2 decimal places.

For propanoic acid, pK_a = 4.87

pH ____________________

(4)

(d) For butanoic acid, $K_a = 1.51 \times 10^{-5}$ mol dm^{-3}

20.0 cm^3 of 0.100 mol dm^{-3} sodium hydroxide solution are added to
25.0 cm^3 of 0.100 mol dm^{-3} butanoic acid solution.

Calculate the pH of the solution formed.

pH ____________________

(5)

(e) A student plans to titrate butanoic acid solution with a solution of ethylamine.

Explain why this titration could **not** be done using an indicator.

____

____

____

____

(2)

(a) Thymol blue

1

(b)

$$K_a = \frac{[H^+][CH_3CH_2COO^-]}{[CH_3CH_2COOH]}$$

Square brackets essential

1

(c) **M1** $K_a = 10^{-pKa} = 1.35 \times 10^{-5}$

M2 $\frac{[H^+]^2}{[CH_3CH_2COOH]}$ OR $[H^+]^2 = K_a\,[CH_3CH_2COOH]$

M2 Square brackets or with numbers

M3 $[H^+] = \sqrt{(1.35 \times 10^{-5} \times 0.1)} = 1.16 \times 10^{-3}$ mol dm^{-3}

M4 pH = $-\log_{10}(1.16 \times 10^{-3}) = 2.94$

M4 = $-\log_{10}$ M3

Answer to 2 decimal places

Allow 2.93

4

(d) **M1** Initial amount of butanoic acid = $25 \times 0.1 \times 10^{-3} = \underline{2.5 \times 10^{-3}}$ mol

M2 Initial amount of NaOH = $20 \times 0.1 \times 10^{-3} = \underline{2.0 \times 10^{-3}}$ mol

M3 Final amount of acid = $2.5 \times 10^{-3} - 2.0 \times 10^{-3} = 5.0 \times 10^{-4}$ mol

***M3** = M1-M2*

M4 $[H^+] = \frac{K_a \times [HX]}{[X^-]}$

Or $[H^+] = \frac{1.51 \times 10^{-5} \times 0.0111}{0.0444}$

***M4** allow volumes cancelled out*

$\frac{1.51 \times 10^{-5} \times 5.0 \times 10^{-4}}{2.0 \times 10^{-3}}$

***M4** Allow $[H^+] = 3.775 \times 10^{-6}$ mol dm^{-3}*

*Alternative method for **M4***

$pH = pKa + \log\frac{[X^-]}{[HX]} = 4.82 + \log\left(\frac{0.0444}{0.0111}\right)$

M5 pH = 5.42

***M5** = dependent on a correct expression for $[H^+]$ in M4*

5

(e) **M1** This is a weak acid and weak base/alkali titration

M2 pH change is too gradual/not sharp (at the equivalence point so colour change of indicator is difficult to judge)

***M2** Allow no vertical/steep section on pH curve*