Q1.

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.

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____

____

(2)

(Total 13 marks)

Q2.

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 / mol² 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 ____

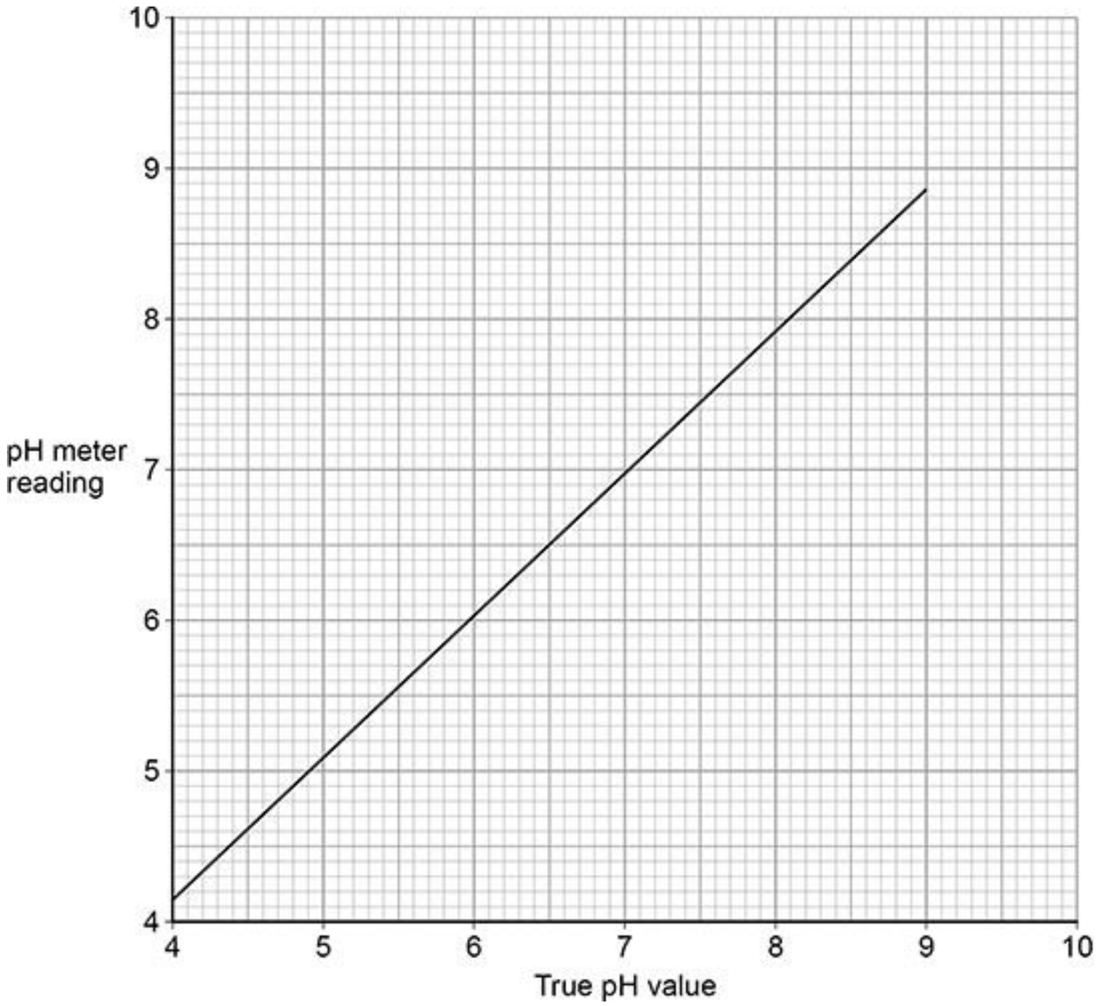
____

(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

____

(1)

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

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____

(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

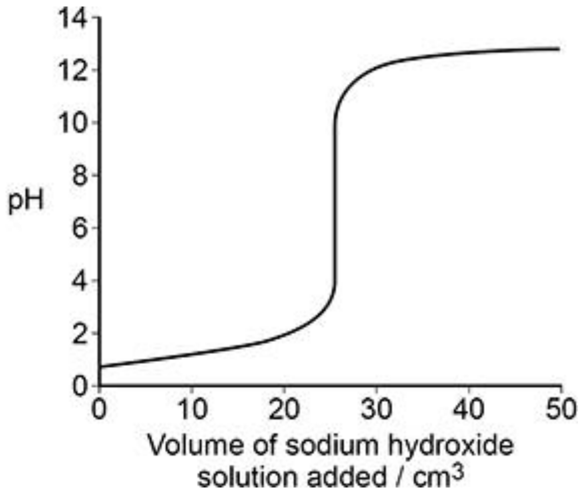


Table 2 shows data about some indicators.

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.

___

___

(1)

(h) 36.25 cm^3 of 0.200 mol dm^{-3} sodium hydroxide solution are added to 25.00 cm^3 of 0.150 mol dm^{-3} hydrochloric acid.

Calculate the pH of the final solution at 25 ºC

$K_w = 1.00 \times 10^{-14}$ mol^2 dm^{-6} at 25 ºC

pH ________________

(5)

(Total 16 marks)

Q3.

Which statement about pH is correct?

A The pH of a weak base is independent of temperature. ☐

B At temperatures above 298 K, the pH of pure water is less than 7. ☐

C The pH of 2.0 mol dm^{-3} nitric acid is approximately 0.30 ☐

D The pH of 0.10 mol dm^{-3} sulfuric acid is greater than that of 0.10 mol dm^{-3} hydrochloric acid. ☐

(Total 1 mark)

Q4.

A 0.10 mol dm^{-3} aqueous solution of an acid is added slowly to 25 cm^3 of a 0.10 mol dm^{-3} aqueous solution of a base.

Which acid–base pair has the highest pH at the equivalence point?

A CH_3COOH and NaOH ☐

B CH_3COOH and NH_3 ☐

C HCl and NaOH ☐

D HCl and NH_3 ☐

(Total 1 mark)

Q5.

This question is about different pH values.

(a) For pure water at 40 °C, pH = 6.67
A student thought that the water was acidic.

Explain why the student was incorrect.

Determine the value of K_w at this temperature.

Explanation ____

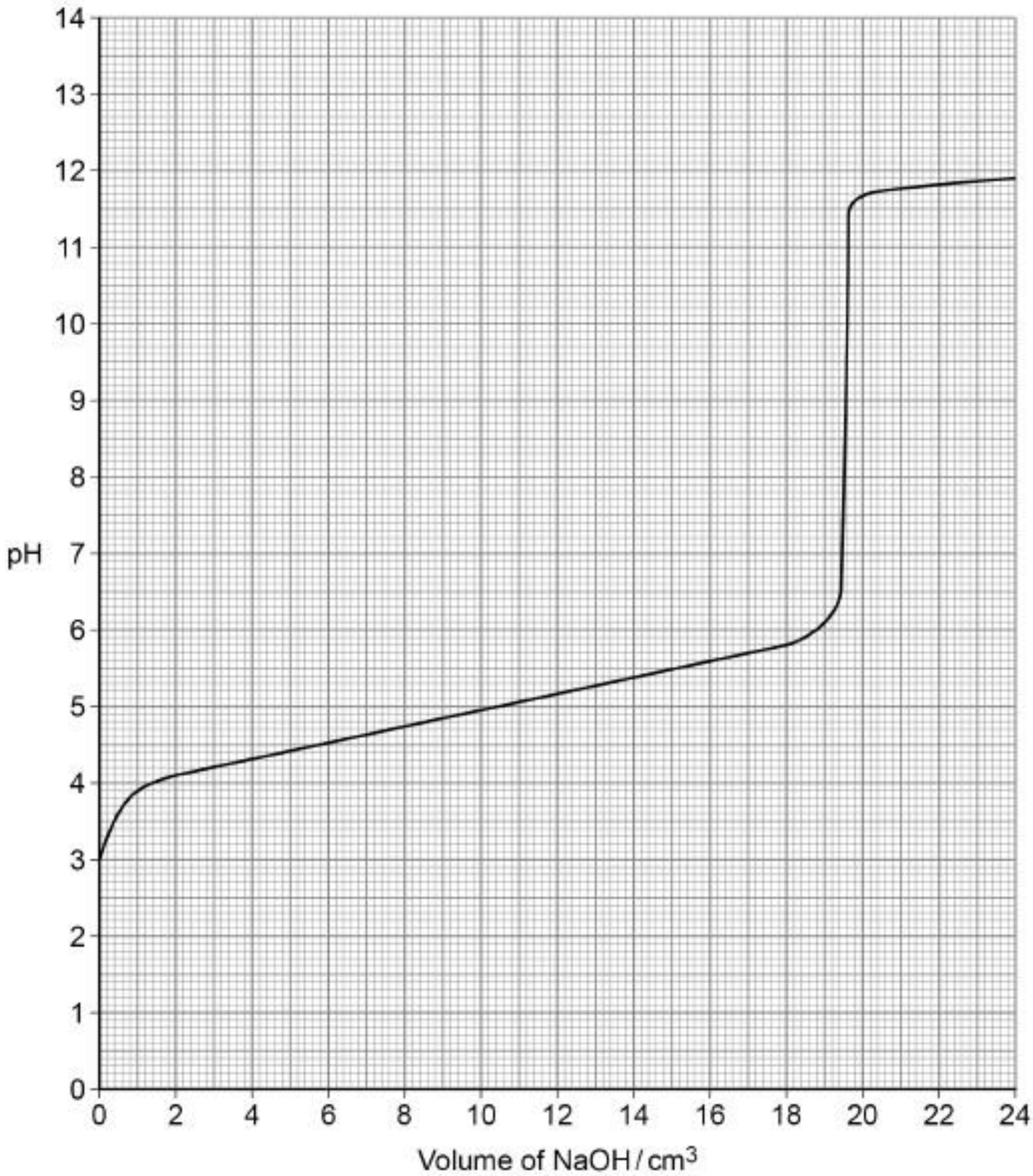
____

K_w ____________________________ $mol^2\ dm^{-6}$

(4)

(b) Sodium hydroxide solution was added gradually from a burette to 25 cm^3 of 0.080 mol dm^{-3} propanoic acid at 25 °C
The pH was measured and recorded at regular intervals.

The results are shown in the diagram.



Use the diagram above to determine the value of K_a for propanoic acid at 25 °C

Show your working.

K_a ________________________________ mol dm^{-3}

(3)

(c) Suggest which indicator is the most appropriate for the reaction in part **(b)**?
Tick (✓) **one** box.

Indicator	**pH range**	**Tick (✓) one box**
methyl orange	3.1 - 4.4	
bromothymol blue	6.0 - 7.6	
cresolphthalein	8.2 - 9.8	
indigo carmine	11.6 - 13.0	

(1)

(d) A student prepared a buffer solution by adding 0.0136 mol of a salt KX to 100 cm^3 of a 0.500 mol dm^{-3} solution of a weak acid HX and mixing thoroughly.

The student then added 3.00×10^{-4} mol of potassium hydroxide to the buffer solution.

Calculate the pH of the buffer solution after adding the potassium hydroxide.

For the weak acid HX at 25 °C the value of the acid dissociation constant, $K_a = 1.41 \times 10^{-5}$ mol dm^{-3}.

Give your answer to two decimal places.

pH ______________________________

(6)

(e) A buffer solution has a constant pH even when diluted.

Use a mathematical expression to explain this.

(1)
(Total 15 marks)

Q6.

Which indicator should be used in a titration to find the concentration of a solution of methylamine using 0.010 mol dm^{-3} hydrochloric acid?

A	Thymol blue	(pH range 1.2–2.8)	⬭
B	Bromophenol blue	(pH range 3.0–4.6)	⬭
C	Phenol red.	(pH range 6.8–8.4)	⬭
D	Phenolphthalein	(pH range 8.3–10.0)	⬭

(Total 1 mark)

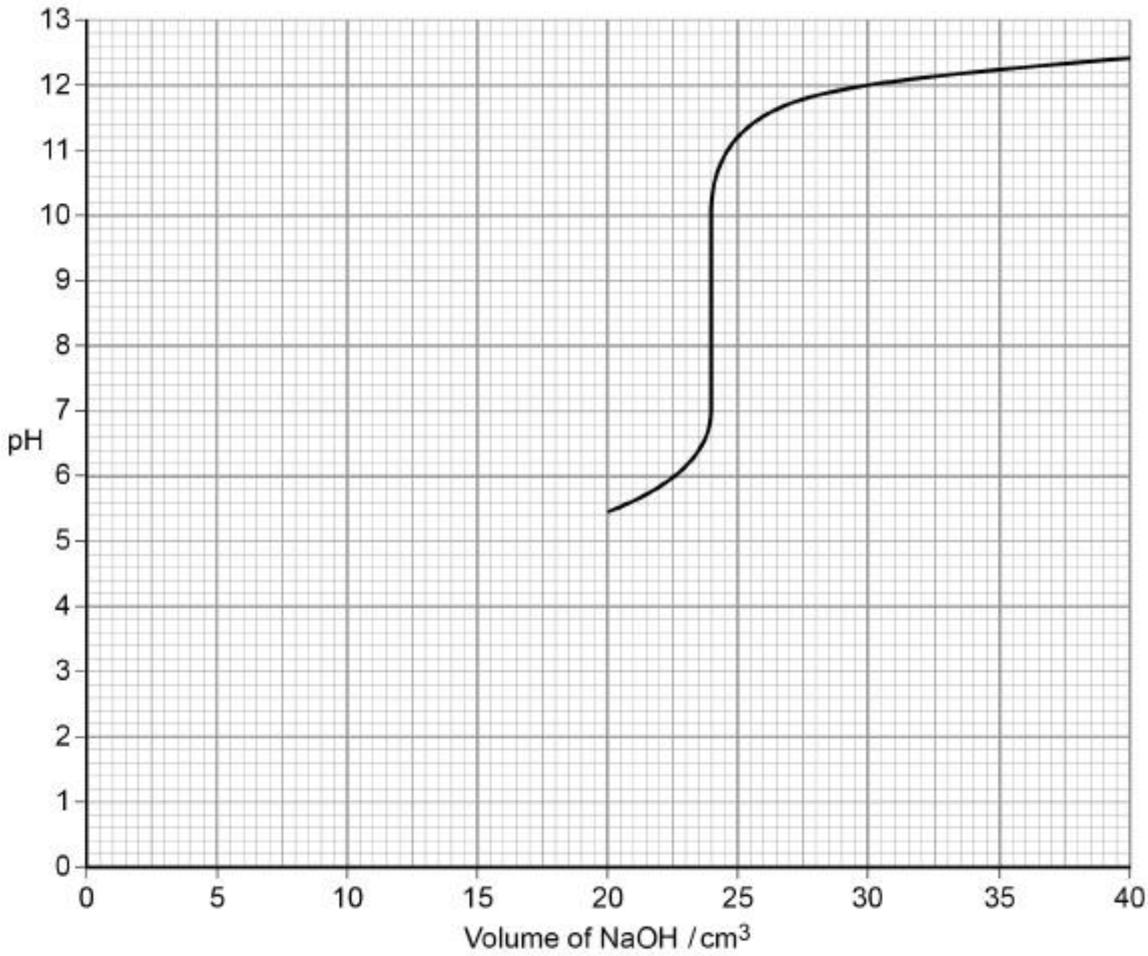
Q7.

A 0.100 mol dm^{-3} solution of sodium hydroxide was gradually added to 25.0 cm^3 of a solution of a weak acid, HX, in the presence of a suitable indicator.

A graph was plotted of pH against the volume of sodium hydroxide solution, as shown in the figure below.

The first pH reading was taken after 20.0 cm^3 of sodium hydroxide solution had been added.

The acid dissociation constant of HX, K_a, = 2.62 × 10^{-5} mol dm^{-3}



(a) The pH range of an indicator is the range over which it changes colour.

Suggest the pH range of a suitable indicator for this titration.

____

(1)

(b) Give the expression for the acid dissociation constant of HX.

K_a =

(1)

(c) Calculate the concentration of HX in the original solution.

Concentration ____________________ mol dm^{-3}

(2)

(d) Calculate the pH of the solution of HX before the addition of any sodium hydroxide.

(If you were unable to calculate a value for the concentration of HX in part (c) you should use a value of 0.600 mol dm^{-3} in this calculation. This is not the correct value.)

pH of HX ____________________

(2)

(e) Calculate the pH of the solution when half of the acid has reacted.

pH of solution ____________________

(1)

(f) Plot your answers to part (d) and part (e) on the grid in the figure above.

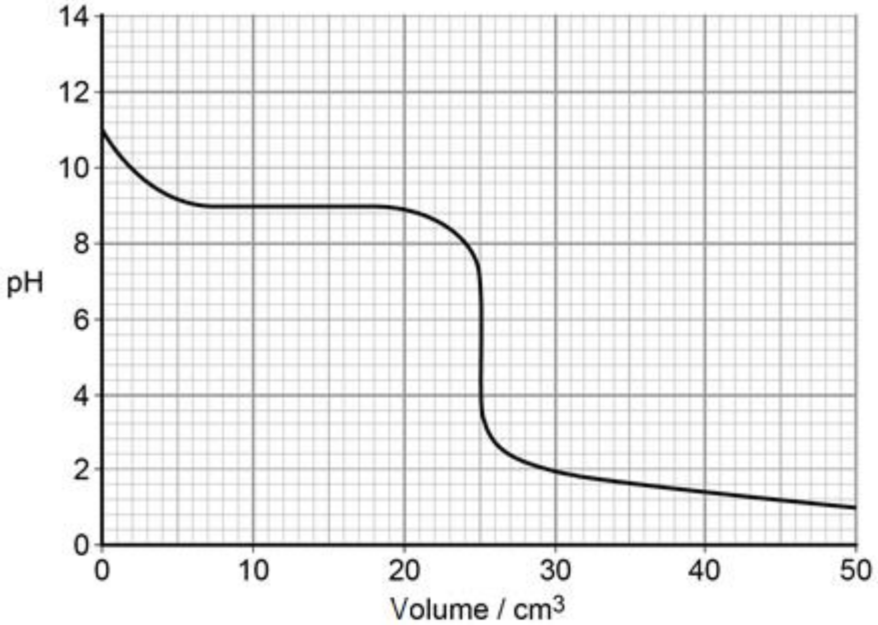
Use these points to sketch the missing part of the curve between 0 and 20 cm^3 of NaOH solution added.

(2)

(Total 9 marks)

Q8.

The graph was obtained from an experiment in which an acid was reacted with an alkali.



(a) Suggest possible formulae for an acid and an alkali that could be used to produce the curve shown in the graph.

Acid ____

Alkali ____

(2)

(b) Suggest briefly a practical procedure that a student could use to obtain data from which the curve in the graph could be plotted.

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____

____

____

____

(3)

(c) The student was provided with samples of three different indicators.

Suggest how the practical procedure in part **(b)** could be refined by the student to identify the most suitable indicator.

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___

___

___

___

(2)

(Total 7 marks)

Q1.

(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*

2

[13]

Q2.

(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 × 0.200 ÷ 1000 = 7.25 × 10^{-3} mol **and** Amount of H^+ = 25.0 × 0.150 ÷ 1000 = 3.75 × 10^{-3} mol

1

M2 Amount of excess OH^- = 7.25 × 10^{-3} – 3.75 × 10^{-3} = 3.50 × 10^{-3} mol

1

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

M3 $[OH^-]$ = (M2) ÷ (61.25 × 10^{-3})

1

M4 $[H^+]$ = 1.00 × 10^{-14} ÷ 5.71 × 10^{-2} = 1.75 × 10^{-13}

M4 $[H^+]$ = 1.00 × 10^{-14} ÷ 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

[16]

Q3.

B

[1]

Q4.

A

[1]

Q5.

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

***M1**: accept equal number/amounts of H^+ and OH^-*

1

M2: $[H^+]$ (= 10^{-pH}) = 2.138×10^{-7}

M2: allow 2.14×10^{-7} 1

M3: $K_w = [H^+]^2$ or $(2.138 \times 10^{-7})^2$

M3: allow (M2)^2 1

M4: $K_w = 4.57 \times 10^{-14}$

M4: allow 4.58×10^{-14}

M4** is dependent on (an answer)*^{2} *in **M3 1

(b) View with Figure X (ie graph) as they may show working there.

Ignore calculations of mols of salt or acid

M1: Determines volume at half equivalence (= $\frac{19.5}{2}$ cm^3) = 9.75 (cm^3)

***M1**: Allow reading on graph to be from 19.4 to 19.7 giving **M1** = 9.7 to 9.85* 1

M2: pH = 4.80 to 4.95

***M2**: Reads off pH at half equivalence* 1

M3: K_a (= 10^{-pH}) = $10^{-4.9}$ = 1.26×10^{-5}

***M3**: Allow* 1.12×10^{-5} *to* 1.58×10^{-5}

***M3**: Allow 2sf or more* 1

Alternative method

M1: pH of pure acid = 3

M2: $K_a = (10^{-3})^2 / 0.080$

M3: = 1.25×10^{-5}

*Alternative **M1** if calculation incorrect:*

Allow pH = pK_a or $[H^+] = K_a$ at <u>half equivalence</u>

(c) cresolphthalein 1

(d)

M1: $K_a = \frac{[H^+][X^-]}{[HX]}$ or $[H^+] = \frac{K_a \times [HX]}{[X^-]}$

***M1**:* allow $[H^+] = \frac{K_a \times [acid]}{[salt]}$ 1

M2: amount of HX = 0.0500 mol 1

M3: amount of HX after add^n of KOH = 0.05 − $\underline{3 \times 10^{-4}}$ = 0.0497 mol

M3: = M2 – 3 × 10^{-4}

1

M4: amount of KX after add^n of KOH = 0.0136 + $\underline{3 \times 10^{-4}}$ = 0.0139 mol

1

M5: $[H^+] = \left(\frac{1.41 \times 10^{-5} \times 0.0497}{0.0139}\right) = \underline{5.04(15) \times 10^{-5}}$

1

M6: pH = $-\log_{10} 5.04(15) \times 10^{-5}$ = 4.30

Answer to 2 decimal places

1

*If no attempt at **M3** and **M4** max 2 marks*

*If **M3** or **M4** attempted using 3 × 10^{-4} max 4 (**M1**, **M2**, **M3** or **M4** and **M6**)*

(e)

ratio $\frac{[HX]}{[X^-]}$

Allow inverse expression

1

[15]

Q6.

B

[1]

Q7.

(a) 7–10.2

any range (i.e. 2 values) within this range

1

(b) $K_a = \frac{[H^+][X^-]}{[HX]}$

ALLOW H_3O^+ for H^+ and A for X

IGNORE $[H^+]^2/[HX]$

must be square brackets

IGNORE state symbols

1

(c) Amount NaOH = (24.0 × 0.100)/1000 = 2.40 × 10^{-3} mol

(= amount HX)

Conc HX = 2.40 × 10^{-3}/0.025 = 0.0960 mol dm^{-3}

ecf for M1/0.025

(d) (K_a = 2.62 × 10^{-5} = $[H^+]^2$/0.0960)

$[H^+]$ = √(2.62 × 10^{-5} × 0.0960) (= 1.59 × 10^{-3} mol dm^{-3})

ecf from part (c) $[H^+]$ = √(2.62 × 10^{-5} × ans to part (c))

From alternative data
$[H^+] = \sqrt{(2.62 \times 10^{-5} \times 0.600)}$ (= 3.96×10^{-3} mol dm^{-3})

(pH = –log 1.59×10^{-3} =) 2.80 (must be 2 or more dp)

1

pH = 2.40 (must be 2 or more dp)
M2 dependent on a calculation of $[H^+]$

1

(e) (pH at half-neutralisation = pK_a)

= –log 2.62×10^{-5} = 4.58 (must be 2 or more dp)
ALLOW 1dp if already penalised in part (d)

1

(f) Both points plotted correctly and line touches both points
ecf from (d) and (e) within 1 small square

1

<u>Line steeper at start</u> then levels (to show buffering)
Mark independently

1

[9]

Q8.

(a) Formula of any strong acid (e.g. HCl)

1

Formula of a weak alkali (e.g. NH_3)

1

(b) Place a fixed volume of alkali in a flask or beaker

1

Add acid in small portions from a burette

1

Stir and use a pH meter to record the pH after each addition of acid

1

(c) Repeat the experiment with each indicator

1

Select the indicator that changes colour rapidly when the pH changes from about 7 to 4

1

[7]