Mark schemes

Q1.

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

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

M3 pH = 6.77

M3 = $-\log_{10}$ 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 $[OH^-] = \frac{3.25\times10^{-3}}{55.0\times10^{-3}} = 0.0591 \text{ mol dm}^{-3}$

M3 = $\frac{M2}{55 \times 10^{-3}}$

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

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

M5 12.31

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

5

[8]

Q2.

(a) Only a (small) fraction of the molecules dissociate (into ions)/ionise (when added to water/in aqueous solution)
OR
partially dissociates/does not fully dissociate (into ions)/ionise (when added to water/in aqueous solution).

Not 'reaction is reversible'

Not if incorrect ions suggested

Not “;Hydrogen ions do not fully dissociate”

1

(b) As there is a large pH change (for a small addition of alkali)

Need idea of rapid/large change in pH

Allow pH change is rapid/gradient is steep (at equivalence point)

Ignore equivalence/end point is very steep

Ignore so they do not miss the equivalence point

1

(c) **M1** $K_a = \frac{[H^+][CH_3CH_2COO^-]}{[CH_3CH_2COOH]}$

Allow [H_3O^+] for [H^+]

NOT () for concentration

NOT $K_a = \frac{[H^+]^2}{[CH_3CH_2COOH]}$

M2 [CH_3CH_2COOH] = [$CH_3CH_2COO^-$]

Allow [HA] = [A^-] / mol/amount of HA = mol/amount of A^-

Allow [HA]/[A^-] OR [A^-]/[HA] = 1

M3 K_a = [H^+] (therefore pH = pK_a)

Allow Log 1 = 0

3

(d) **M1** Allow pH between 4.4 - 4.7

M2 Answer from $K_a = 10^{-M1}$ to min 2sf

M2 ECF from M1 only if M1 is in range 4.0 to <4.4 or >4.7 to 5.0

If M1 in range shown, M2 = 1.99526×10^{-5} to 3.9811×10^{-5} to at least 2 sf

Using pH at start (= 2.80) scores 0/2

2

(e) **M1** OH^- reacts with propanoic acid OR reacts with H^+

M2 EITHER

ratio of $[CH_3CH_2COOH]$ to $[CH_3CH_2COO^-]$ remains almost constant

OR

Equilibrium **for dissociation of CH_3CH_2COOH** moves right to maintain $[H^+]$ /replace H^+

Allow CH_3CH_2COOH dissociates to maintain $[H^+]$ /replace H^+

2

(f) **M1** Methyl orange - would not change colour at the equivalence point (allow end-point)/pH range does not match rapid pH change/pH range does not include/match equivalence point

Allow methyl orange changes colour before/below equivalence point

Allow pH range does not fall in the range of pH from 6-11

Ignore reference to this being a weak acid - strong base titration

M2 Universal indicator - idea of range of colours during titration/no distinct colour change (at equivalence/end-point)

2

[11]

Q3.

(a) Thymol blue

1

(b)

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

Square brackets essential

1

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

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]

Q4.

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

amount $Ca(OH)_2$ $\left(=\frac{0.6}{74.1}\right)$ = 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 x 10^{-15} ÷ **M2** (= 6.30 x 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

[7]

Q5.

(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** $[OH^-]$ = **2 ×** 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

1

(e) **M1** (at start) n(NaOH) = $\frac{2.00}{40}$ = <u>0.05</u> (mol)
and
$n(CH_3COOH) = \frac{500 \times 1.0}{1000}$ = <u>0.5</u> (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) = $\frac{2.00}{40}$ = <u>0.05</u> (mol)
and
$n(CH_3COOH) = \frac{500 \times 1.0}{1000}$ = <u>0.5</u> (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

M5 pH = 3.80

answer to 2 decimal places

M5 allow 3.81

1

[16]