OCR A Chemistry A-Level

Module 5 - Physical Chemistry & Transition Elements

Acids and Bases

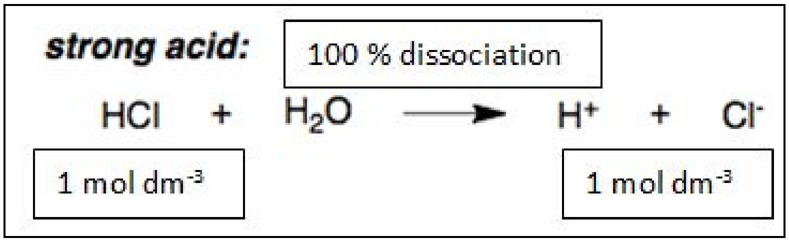
Notes and Example Calculations

Answers given at the end of the booklet

pH Acid Calculations

Strong Acids

To calculate the pH of acids you need to know the concentration of H^+ ions that have been dissociated. In strong acids this is straightforward because they completely dissociate. This means the concentration of the acid = concentration of H^+.



Once $[H^+]$ is calculated the pH can be found out by using the following equation:

$$\boxed{pH = -\log_{10}[H^+]}$$

If given the pH of the acid in solution and need to work out the concentration of the acid use this equation:

$$\boxed{[H^+] = 10^{-pH}}$$

Example 1:

Work out the pH of 0.5 mol dm^{-3} HNO_3.

Step 1: Write the equation to work out the pH

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

Step 2: Input the concentration of the acid into the equation.

$pH = -\log (0.5)$

$= 0.3$

Example 2:

What is the $[H^+]$ ions in HCl, pH 4.3.

Step 1: Use the correct form of the equation.

$[H^+] = 10^{-pH}$

Step 2: Input the pH value to find $[H^+]$.

$[H^+] = 10^{-4.3}$
$= \underline{\mathbf{5.01 \times 10^{-5}}}$

Weak Acids

Calculating the pH of weak acids is a bit more difficult as weak acids don't fully dissociate in solution. To calculate the pH of a weak acids you have to use the acid dissociation constant, Ka.

Example 3:

Write an expression for the acid dissociation constant, Ka for ethanoic acid:

$$CH3COOH \rightleftharpoons CH3COO^- + H^+$$

[Just like Kc , and Kp expressions it is always products over reactants.]

$$Ka = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

Example 4:

Calculate the pH of a 0.025 mol dm^{-3} solution of methanoic acid. For HCOOH, Ka = 1.58×10^{-4} mol dm^{-3}

$$HCOOH \rightleftharpoons HCOO^- + H^+$$

Step 1: Write the Ka expression for HCOOH.

$$Ka = \frac{[HCOO^-][H^+]}{[HCOOH]}$$

Step 2: Rearrange the expression so that $[H^+]$ is the subject.

[The assumption here is that $[H^+] = [HCOO^-]$]

⇒ $[H^+] = \sqrt{(Ka \times [HCOOH])}$

Step 3: Input the values into the equation to work out $[H^+]$.

$[H^+] = \sqrt{(Ka \times [HCOOH])}$
$= \sqrt{(1.58 \times 10^{-4} \times 0.025)}$
$= 1.987 \times 10^{-3}$

Step 4: Input the concentration of the H^+ ions into the pH equation.

⇒ $pH = -\log [H^+]$
$= -\log(1.987 \times 10^{-3})$
= **2.70**

pKa is sometimes used instead of Ka to make numbers more manageable.

$$pKa = -\log(Ka) \qquad Ka = 10^{-pKa}$$

Worked Exam Style Questions

Question 1

A solution of phenol in water has a concentration of 4.7 g dm^{-3}.

(i) Write an expression for the acid dissociation constant, K_a, of phenol.

$C_6H_5OH(aq) \rightleftharpoons H^+(aq) + C_6H_5O^-(aq) \qquad K_a = 1.3 \times 10^{-10}\ mol\ dm^{-3}$

$$\Rightarrow Ka = \frac{[H^+][C_6H_5O^-]}{[C_6H_5OH]}$$

(ii) Calculate the pH of this solution of phenol.

Step 1: Rearrange the Ka expression so that $[H^+]$ is the subject.

$$Ka = \frac{[H^+]^2}{[C_6H_5OH]}$$

⇒ $[H^+] = \sqrt{(Ka \times [C_6H_5OH])}$

Step 2: Change the concentration of phenol from gdm^{-3} to $moldm^{-3}$.

[Divide by molecular mass (remember: mass = Mr x mol)]

Mr of phenol:
C x 6 - 12 x 6 = 72
H x 6 - 6 x 1 = 6
O x 1 - 16 x 1 = 16

⇒ 72 + 6 + 16 = 94

⇒ 4.7 / 94 = 0.05 mol dm^{-3}

Step 3: Input the values into the expression to work out $[H^+]$.

$[H^+] = \sqrt{(Ka \times [C6H5OH])}$
$= \sqrt{(1.3 \times 10^{-10} \times 0.05)}$
$= 2.55 \times 10^{-6}$

Step 4: Input the concentration of the H^+ ions into the pH equation.

⇒ $pH = -\log [H^+]$
$= -\log(2.55 \times 10^{-6})$
= **<u>5.59</u>**

Question 2

Write the expression for K_a for methanoic acid.

$$\Rightarrow \mathbf{Ka = \frac{[H^+][HCOO^-]}{[HCOOH]}}$$

A 1.50×10^{-2} mol dm^{-3} solution of HCOOH has $[H^+] = 1.55 \times 10^{-3}$ mol dm^{-3}.

Calculate the values of K_a and pK_a for methanoic acid.

Step 1: Sub in the values into the Ka expression.

$$\Rightarrow Ka = \frac{[H^+]^2}{[HCOOH]}$$

$= (1.55 \times 10^{-3})^2 / 1.5 \times 10^{-2}$

$= \underline{\mathbf{1.6 \times 10^{-4}}}$

Step 2: Use this value to calculate pKa.

$\Rightarrow$ pKa = -log(Ka)

$= -\log(1.6 \times 10^{-4})$

$= \underline{\mathbf{3.80}}$

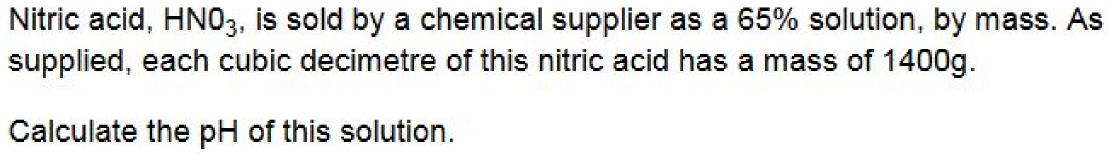
Estimate the percentage of HCOOH molecules that have dissociated in this aqueous solution of methanoic acid.

[Percentage dissociation = ([H$^+$] / concentration of the acid) x 100]

$$\Rightarrow \frac{1.55 \times 10^{-3}}{0.015} \times 100 = \underline{\mathbf{10.3\ \%}}$$

Try these questions…

1.

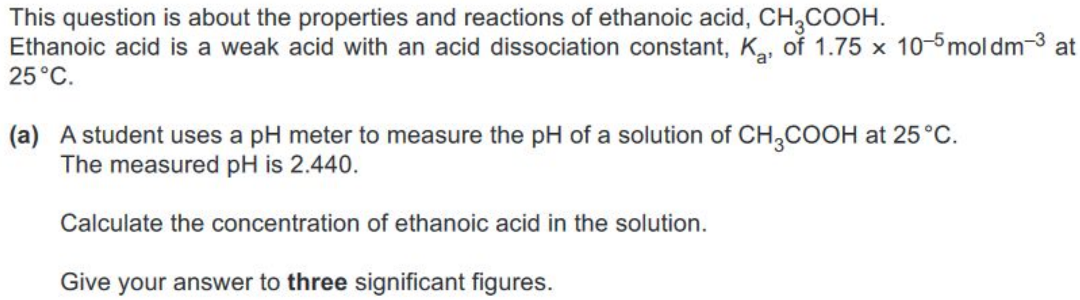


Nitric acid, $HN0_3$, is sold by a chemical supplier as a 65% solution, by mass. As supplied, each cubic decimetre of this nitric acid has a mass of 1400g.

Calculate the pH of this solution.

[3 marks]

2.



This question is about the properties and reactions of ethanoic acid, CH_3COOH.
Ethanoic acid is a weak acid with an acid dissociation constant, K_a, of 1.75×10^{-5} mol dm^{-3} at 25 °C.

(a) A student uses a pH meter to measure the pH of a solution of CH_3COOH at 25 °C. The measured pH is 2.440.

Calculate the concentration of ethanoic acid in the solution.

Give your answer to **three** significant figures.

[3 marks]

pH Base Calculations

Strong Bases

Despite being covalent, water conducts electricity to a very small extent. This is due to the slight ionisation:

$$H2O \rightleftharpoons H^+ + OH^-$$

If we write the Kc expression for this dissociation it would be like this:

$$Kc = \frac{[H^+][OH^-]}{[H2O]}$$

As the dissociation is small, the water concentration is very large compared with the dissociated ions and any changes to its value are insignificant, its concentration can be regarded as constant

This new constant is Kw and is equal to $[H^+][OH^-]$. At room temperature Kw has a fixed value:

$$\mathbf{Kw = 1 \times 10^{-14}\ mol^2\ dm^{-6}}$$

Kw which is called the **ionic product of water** is used to calculate the pH of an alkali.

Example 1:
Work out the pH of 0.05 mol dm^{-3}.

Step 1: Rearrange the Kw expression so that $[H^+]$ is subject.

$Kw = [H^+][OH^-] = 1 \times 10^{-14}$

$$\Rightarrow [H^+] = \frac{Kw}{[OH^-]}$$

Step 2: Input values into the equation.

$$\Rightarrow [H^+] = \frac{1 \times 10^{-14}}{0.05}$$

$$= 2 \times 10^{-13}\ mol\ dm^{-3}$$

Step 3: Input $[H^+]$ to the pH equation.

$$\begin{aligned} \Rightarrow pH &= -\log [H^+] \\ &= -\log (\,2 \times 10^{-13}) \\ &= \underline{\mathbf{12.7}} \end{aligned}$$

Example 2:
Calculate the pH of 0.0450 mol dm^{-3} $Ba(OH)_2$.

Step 1: Rearrange the Kw expression so that $[H^+]$ is subject.

$Kw = [H^+][OH^-] = 1 \times 10^{-14}$

$$\Rightarrow [H^+] = \frac{Kw}{[OH^-]}$$

Step 2: Input values into the equation.

$$\Rightarrow [H^+] = \frac{1 \times 10^{-14}}{(0.045 \times 2)}$$

[The concentration of the OH^- ions is multiplied by 2 because the ratio between barium ions and hydroxide ions is 1:2]

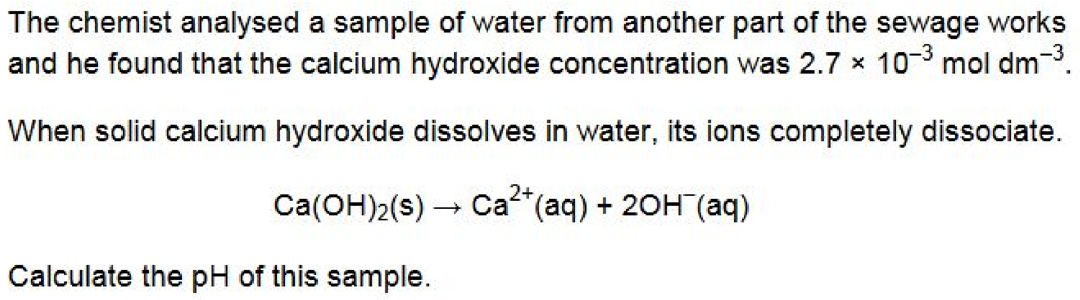
$\Rightarrow 1.1 \times 10^{-13}$ mol dm^{-3}

Step 3: Input $[H^+]$ to the pH equation.

$\Rightarrow$ pH = -log $[H^+]$
= -log (1.1×10^{-13})
= **<u>12.95</u>**

Worked Exam Style Questions

Question 1



The chemist analysed a sample of water from another part of the sewage works and he found that the calcium hydroxide concentration was 2.7×10^{-3} mol dm^{-3}.

When solid calcium hydroxide dissolves in water, its ions completely dissociate.

$$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$$

Calculate the pH of this sample.

Step 1: Rearrange the Kw expression so that $[H^+]$ is subject.

$Kw = [H^+][OH^-] = 1 \times 10^{-14}$

$$\Rightarrow [H^+] = \frac{Kw}{[OH^-]}$$

Step 2: Input values into the equation.

$$\Rightarrow [H^+] = \frac{1 \times 10^{-14}}{(2.7 \times 10^{-3} \times 2)}$$

[The concentration of the OH^- ions is multiplied by 2 because the ratio between calcium ions and hydroxide ions is 1:2]

$\Rightarrow 1.85 \times 10^{-12}$ mol dm^{-3}

Step 3: Input $[H^+]$ to the pH equation.

$\Rightarrow pH = -\log [H^+]$
$= -\log (1.85 \times 10^{-12})$
= **11.73**

Try these questions…

3. A student measured the pH of a solution of sodium hydroxide as 13.54 at 25 °C.

$K_w = 1.0 \times 10^{-14}$ mol^2 dm^{-6} at 25 °C.

(i) Write down an expression for the ionic product, K_w, for water.

(ii) Calculate the concentration, in mol dm^{-3}, of this solution of sodium hydroxide.

[3 marks]

4. Water dissociates slightly according to the equation:

$$H_2O(l) \text{ ----> } H^+(aq) + OH^-(aq)$$

The ionic product of water, K_w, is given by the expression

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

K_w varies with temperature as shown in the table.

Temperature / °C	**K_w / mol^2 dm^{-6}**
25	1.00×10^{-14}
50	5.48×10^{-14}

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

[3 marks]

Buffers

A buffer solution minimises changes in pH during the addition of small amounts of acid or alkali. It is a mixture of weak acid and its conjugate base.

For example in a CH3COOH/CH3COONa buffer system CH3COOH is the weak acid and the conjugate base is $CH3COO^-$.

The pH of a buffer solution is affected by the acid dissociation constant Ka and the concentration ratio of the weak acid and its conjugate base.

For a buffer consisting of a weak acid, CH3COOH and its conjugate base, $CH3COO^-$:

$$Ka = \frac{[CH3COO^-][H^+]}{[CH3COOH]}$$

Example 1:

pH of buffer solution made from excess weak acid and strong base.

50 cm^3 of 0.250 mol dm^{-3} butanoic acid added to 50 cm^3 of 0.05 mol dm^{-3} sodium hydroxide.

Ka = 1.51×10^{-5}. What is the pH?

Step 1: Write balanced equation.

⇒ $NaOH + C_4H_7COOH \rightarrow C_4H_7COONa + H_2O$

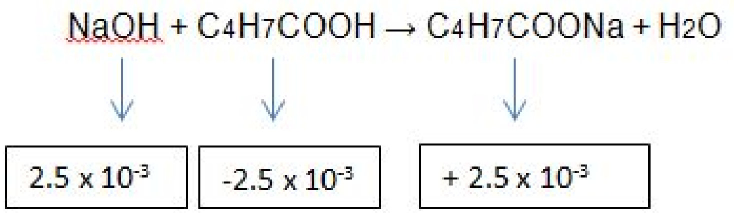
Step 2: Work out the moles of NaOH (base).

[Moles = concentration x volume]

⇒ $0.05 \times (50/1000) = 2.5 \times 10^{-3}$

Step 3: State how many moles of acid would react with this many moles of base and how many moles of salt would be formed

⇒



Step 4: Work out the moles of acid left and concentration of the buffer solution.

⇒ Moles of CH_3COOH at the start = concentration x volume
= (50/1000) x 0.25
= 0.0125

Moles of acid left after reaction = $0.0125 - 2.5 \times 10^{-3}$
= 0.01

Concentration of CH_3COOH = moles / volume (dm^3)
= 0.01 / 0.1
= 0.1 mol dm^{-3}

Step 5: Work out the concentration of the conjugate base CH_3COO^-.

⇒ Concentration of CH_3COO^- = moles / volume
= $2.5 \times 10^{-3} / 0.1$
= 0.025 mol dm^{-3}

Step 6: Work out [H+].

$$Ka = \frac{[CH3COO^-][H^+]}{[CH3COOH]}$$

$$\Rightarrow [H^+] = \frac{Ka \times [CH3COOH]}{[CH3COO^-]}$$

$$= \frac{1.51 \times 10^{-5} \times (0.1)}{0.025}$$

$$= 6.04 \times 10^{-5}$$

Step 7: Work out the pH of the buffer solution by inputting $[H^+]$ into the pH equation.

$$\begin{aligned} pH &= -\log[H^+] \\ &= -\log(6.04 \times 10^{-5}) \\ &= \underline{\mathbf{4.22}} \end{aligned}$$

Example 2:

Calculate the pH of a buffer solution which contains the weak monoprotic acid, propanoic acid (CH_3CH_2COOH), in concentration 0.1 mol dm^{-3} and sodium propanoate in concentration 0.05 mol dm^{-3}. K_a of propanoic acid is 1.26×10^{-5} mol dm^{-3}.

Step 1: Write the acid dissociation constant Ka expression.

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

Step 2: Rearrange the Ka expression so that [H+] is the subject.

$$\Rightarrow [H^+] = \frac{Ka \times [CH_3CH_2COOH]}{[CH_3CH_2COO^-]}$$

Step 3: Input the values to work out $[H^+]$.

$$\Rightarrow [H^+] = \frac{1.26 \times 10^{-5} \times (0.1)}{(0.05)}$$

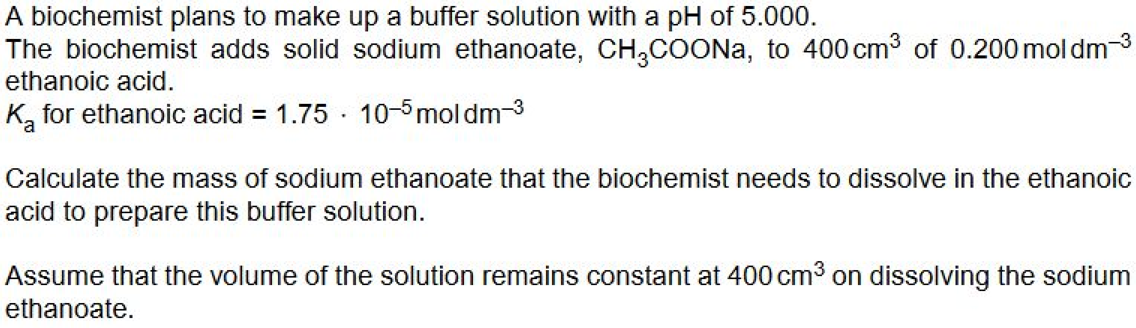
$$= 2.52 \times 10^{-5} \text{ mol dm}^{-3}$$

Step 4: Work out the pH of the buffer solution.

$$\Rightarrow \text{pH} = -\log[H^+]$$
$$= -\log(2.52 \times 10^{-5})$$
$$= \underline{\mathbf{4.60}}$$

Worked Exam Style Questions

Question 1



A biochemist plans to make up a buffer solution with a pH of 5.000.
The biochemist adds solid sodium ethanoate, CH_3COONa, to 400 cm^3 of 0.200 mol dm^{-3} ethanoic acid.
K_a for ethanoic acid = 1.75 · 10^{-5} mol dm^{-3}

Calculate the mass of sodium ethanoate that the biochemist needs to dissolve in the ethanoic acid to prepare this buffer solution.

Assume that the volume of the solution remains constant at 400 cm^3 on dissolving the sodium ethanoate.

Step 1: Work out $[H^+]$ of the buffer solution with pH of 5.000.

$$\Rightarrow [H^+] = 10^{-pH}$$
$$= 10^{-5}$$
$$= 1 \times 10^{-5} \text{ mol dm}^{-3}$$

Step 2: Write the acid dissociation constant Ka expression.

$$\Rightarrow Ka = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

Step 3: Work out the concentration of the conjugate base by inputting the values for the corresponding substances.

$$\Rightarrow [CH3COO^-] = \frac{Ka \times [CH3COOH]}{[H^+]}$$
$$= \frac{1.75 \times 10^{-5} \times (0.2)}{1 \times 10^{-5}}$$

$$= 0.35 \text{ mol dm}^{-3}$$

Step 4: Work out the number of moles the conjugate base (sodium ethanoate).

⇒ Moles = concentration x volume
= 0.35 x (400/1000)
= 0.14 mol

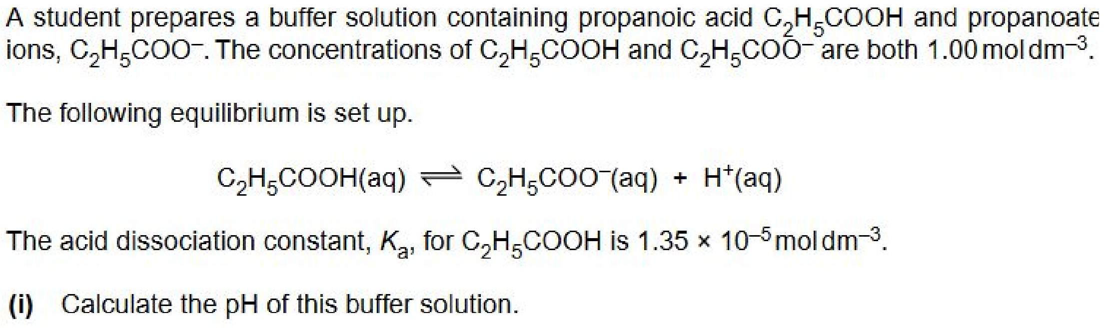
Step 5: Work out the mass of sodium ethanoate.

[Mass = Mr x mol]

⇒ Mr of CH_3COONa = (12 x 2) + (16 x 2) + 3 + 23 = 82

⇒ 82 x 0.14 = **11.48 g**

Question 2



A student prepares a buffer solution containing propanoic acid C_2H_5COOH and propanoate ions, $C_2H_5COO^-$. The concentrations of C_2H_5COOH and $C_2H_5COO^-$ are both 1.00 mol dm^{-3}.

The following equilibrium is set up.

$$C_2H_5COOH(aq) \rightleftharpoons C_2H_5COO^-(aq) + H^+(aq)$$

The acid dissociation constant, K_a, for C_2H_5COOH is 1.35×10^{-5} mol dm^{-3}.

(i) Calculate the pH of this buffer solution.

[1 mark]

Step 1: Calculate the concentration of H^+ ions.

$$\Rightarrow [H^+] = \frac{Ka \times [C_2H_5COOH]}{[C_2H_5COO^-]}$$

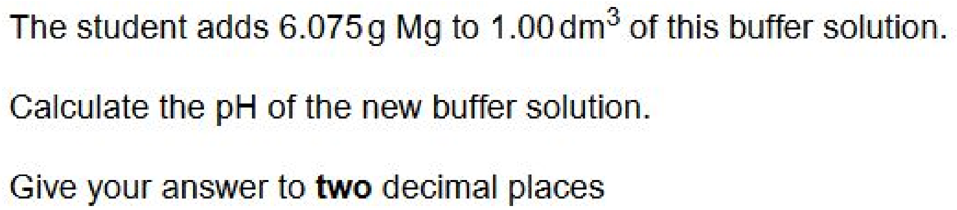
$$= \frac{1.35 \times 10^{-5} \times 1}{1}$$

$$= 1.35 \times 10^{-5}$$

Step 2: Work out the pH of the buffer solution.

⇒ pH = -log(1.35×10^{-5})
= **4.87**

Question 3



The student adds 6.075 g Mg to 1.00 dm^3 of this buffer solution.

Calculate the pH of the new buffer solution.

Give your answer to **two** decimal places

[5 marks]

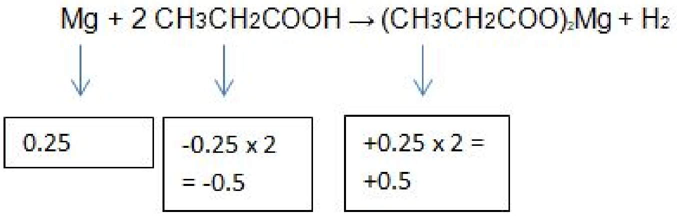
Step 1: Write a balanced equation of this reaction.

⇒ $Mg + 2CH_3CH_2COOH \rightarrow (CH_3CH_2COO)_2Mg + H_2$

Step 2: Work out the moles of magnesium added.

⇒ Moles = mass / Mr
= 6.075 / 24.3
= 0.25

Step 3: Work out how many moles of will react and how many moles of the salt formed using stoichiometry.



Step 4: Work out how many moles of acid/salt after the reaction:

⇒ Moles of CH_3CH_2COOH and $CH_3CH_2COO^-$ = 1.00.

$$Mg + 2 CH3CH2COOH \rightarrow (CH3CH2COO)_2Mg + H2$$

1 - (2 x 0.25) = 0.5 mol   1 + (2 x 0.25) = 1.5 mol

For the acid you subtract 0.5 because 0.5 moles react so are used up.
For the salt 0.5 moles is added because this is the product that is made.

Step 4: Work out the concentration of the acid and conjugate base.

[CH3CH2COOH] = 0.5 / 1 = 0.5 mol dm^{-3}
[CH3CH2COO^-] = 1.5 / 1 = 1.5 mol dm^{-3}

Step 5: Write and rearrange the acid dissociation constant Ka expression for this reaction.

$$\Rightarrow [H^+] = \frac{Ka \times [C2H5COOH]}{[C2H5COO^-]}$$

$$= \frac{1.35 \times 10^{-5} \times (0.5)}{1.5}$$

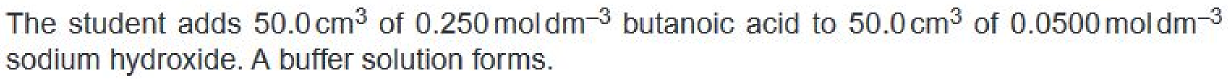
$$= 4.5 \times 10^{-6} \text{ mol dm}^{-3}$$

Step 6: Work of the pH of the buffer solution using the pH equation.

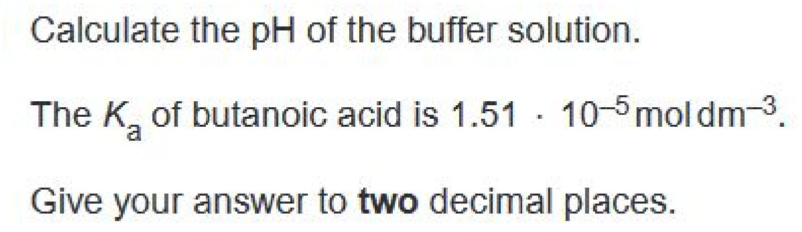
$\Rightarrow$ pH = -log[H^+]
= -log (4.5×10^{-6})
= **5.35**

Try these questions …

5.



The student adds 50.0 cm^3 of 0.250 mol dm^{-3} butanoic acid to 50.0 cm^3 of 0.0500 mol dm^{-3} sodium hydroxide. A buffer solution forms.



Calculate the pH of the buffer solution.

The K_a of butanoic acid is $1.51 \cdot 10^{-5}$ mol dm^{-3}.

Give your answer to **two** decimal places.

[5 marks]

6.

The student plans to prepare a buffer solution that has a pH of 4.50. The buffer solution will contain ethanoic acid, CH_3COOH, and sodium ethanoate, CH_3COONa.

The student plans to add 9.08 g CH_3COONa to 250 cm^3 of 0.800 mol dm^{-3} CH_3COOH. The student assumes that the volume of the solution does not change.

(i) Show by calculation whether, or not, the student's experimental method would produce the required pH.

Show **all** your working.

[5 marks]

Answers

Q1.

(i) $K_w = [H^+(aq)]\,[OH^-(aq)]$ **(1)** 1

state symbols not needed

(ii) $[H^+(aq)] = 10^{-pH} = 10^{-13.54} = 2.88/2.9 \times 10^{-14}$ mol dm^{-3} **(1)**

$$[NaOH] / [OH^-(aq)] = \frac{K_w}{[H^+(aq)]} = \frac{1.0\times10^{-14}}{2.88\times10^{-14}}$$

$= 0.347 / 0.35$ mol dm^{-3} **(1)** 2

[3]

Q2

Question			Answer	Marks	Guidance
21	(a)		**FIRST, CHECK THE ANSWER ON ANSWER LINE** **IF** answer = **0.753**, award **3 marks** --- $[H^+] = 10^{-pH} = 10^{-2.440} = 3.63 \times 10^{-3}$ (mol dm^{-3}) ✓ $[CH_3COOH] = \frac{[H^+]^2}{K_a}$ **OR** $\frac{(3.63 \times 10^{-3})^2}{1.75 \times 10^{-5}}$ ✓ = 0.753 (mol dm^{-3}) ✓	3	**ALLOW** use of **HA** and **A$^-$** **ALLOW 3 SF** up to calculator value of $3.630780548 \times 10^{-3}$ correctly rounded **NOTE:** Answer is same from unrounded $[H^+]$ calculator value and 3 SF $[H^+]$ value **ALLOW** 0.749 if $[H^+]$ has been subtracted from $[CH_3COOH]$ for greater accuracy at end

Q3

(i) $K_w = [H^+(aq)]\,[OH^-(aq)]$ **(1)** 1

state symbols not needed

(ii) $[H^+(aq)] = 10^{-pH} = 10^{-13.54} = 2.88/2.9 \times 10^{-14}$ mol dm^{-3} **(1)**

$[NaOH] / [OH^-(aq)] = \frac{K_w}{[H^+(aq)]} = \frac{1.0 \times 10^{-14}}{2.88 \times 10^{-14}}$

= 0.347 / 0.35 mol dm^{-3} **(1)** 2

[3]

Q4. $[H^+]$ = $\sqrt{K_w}$ (or = $\sqrt{5.48 \times 10^{-14}}$)

Correct pH answer scores 3

1

If wrong method no marks

Using alternative K_w (1.00×10^{-14}) gives pH = 7.<u>00</u> which scores 1

= 2.34×10^{-7}

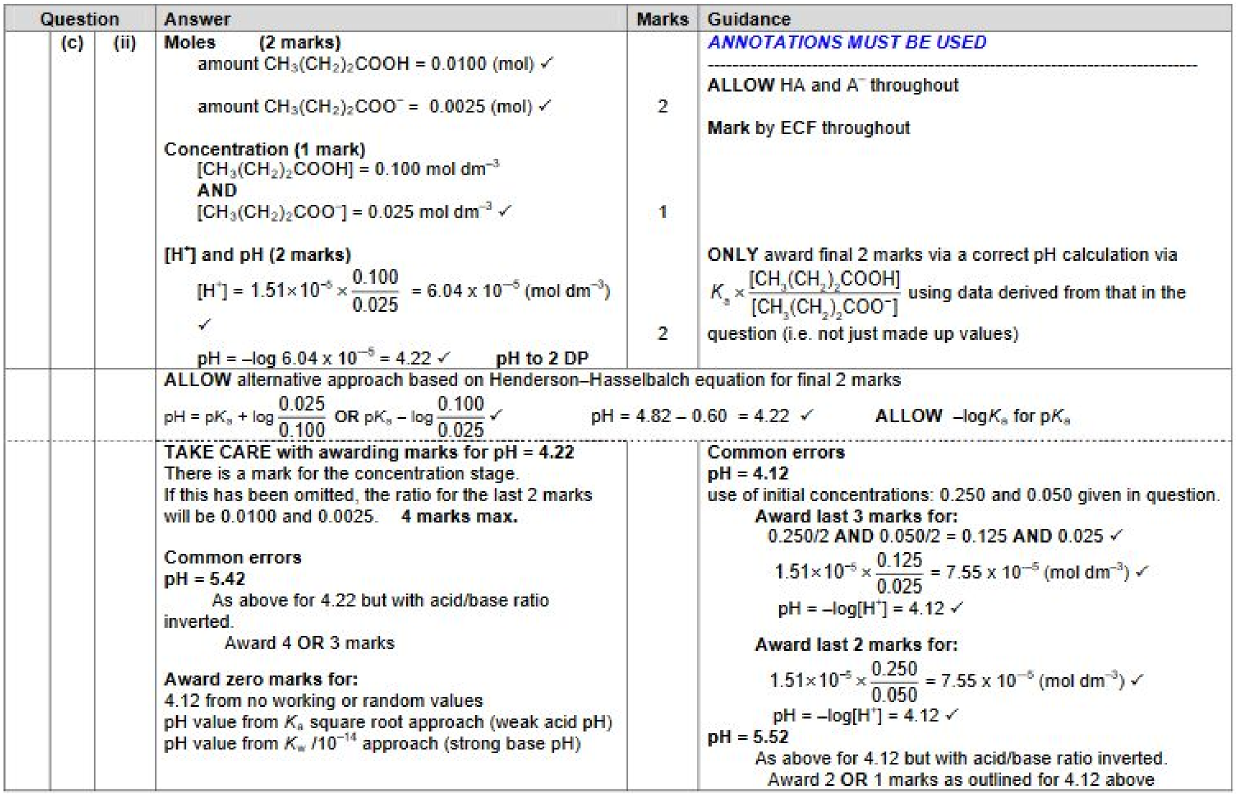
1

pH = 6.63

Final answer must have 2dp

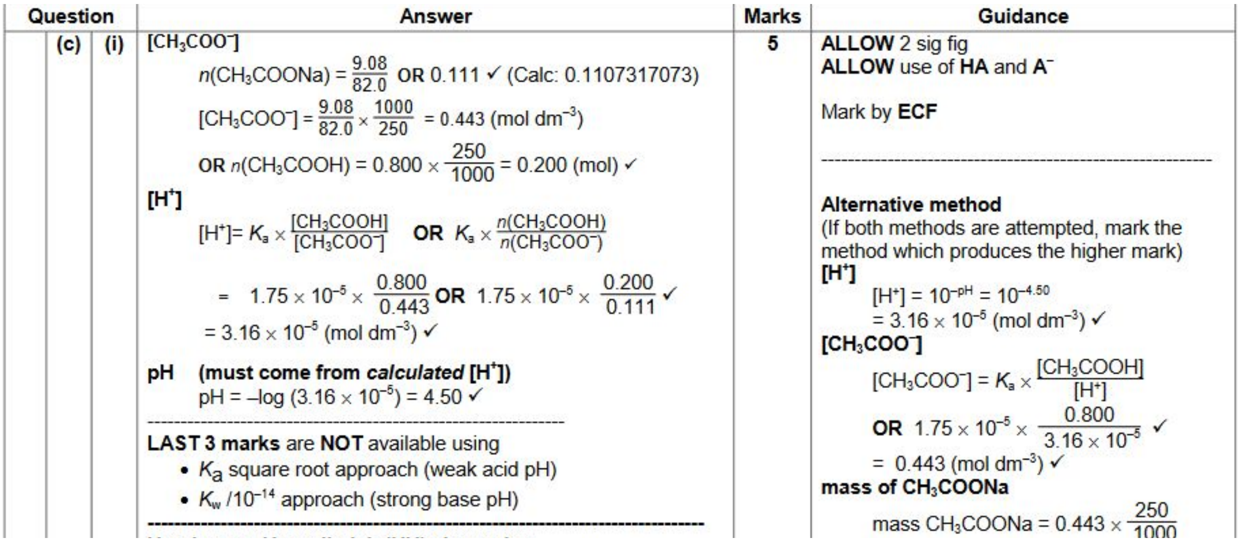
1

Q5



Question			Answer	Marks	Guidance
	(c)	(ii)	**Moles** (2 marks) amount $CH_3(CH_2)_2COOH$ = 0.0100 (mol) ✓ amount $CH_3(CH_2)_2COO^-$ = 0.0025 (mol) ✓	2	***ANNOTATIONS MUST BE USED*** -- **ALLOW** HA and A^- throughout **Mark by ECF throughout**
			Concentration (1 mark) $[CH_3(CH_2)_2COOH]$ = 0.100 mol dm^{-3} **AND** $[CH_3(CH_2)_2COO^-]$ = 0.025 mol dm^{-3} ✓	1	
			[H⁺] and pH (2 marks) $[H^+] = 1.51\times10^{-5}\times\frac{0.100}{0.025}$ = 6.04 x 10^{-5} (mol dm^{-3}) ✓ pH = –log 6.04 x 10^{-5} = 4.22 ✓ **pH to 2 DP**	2	**ONLY** award final 2 marks via a correct pH calculation via $K_a\times\frac{[CH_3(CH_2)_2COOH]}{[CH_3(CH_2)_2COO^-]}$ using data derived from that in the question (i.e. not just made up values)
			ALLOW alternative approach based on Henderson–Hasselbalch equation for final 2 marks $pH = pK_a + \log\frac{0.025}{0.100}$ **OR** $pK_a - \log\frac{0.100}{0.025}$ ✓ pH = 4.82 – 0.60 = 4.22 ✓ **ALLOW** $-\log K_a$ for pK_a		
			TAKE CARE with awarding marks for pH = 4.22 There is a mark for the concentration stage. If this has been omitted, the ratio for the last 2 marks will be 0.0100 and 0.0025. **4 marks max.** **Common errors** **pH = 5.42** As above for 4.22 but with acid/base ratio inverted. Award 4 OR 3 marks **Award zero marks for:** 4.12 from no working or random values pH value from K_a square root approach (weak acid pH) pH value from K_w /10^{-14} approach (strong base pH)		**Common errors** **pH = 4.12** use of initial concentrations: 0.250 and 0.050 given in question. **Award last 3 marks for:** 0.250/2 **AND** 0.050/2 = 0.125 **AND** 0.025 ✓ $1.51\times10^{-5}\times\frac{0.125}{0.025}$ = 7.55 x 10^{-5} (mol dm^{-3}) ✓ pH = –log[H⁺] = 4.12 ✓ **Award last 2 marks for:** $1.51\times10^{-5}\times\frac{0.250}{0.050}$ = 7.55 x 10^{-5} (mol dm^{-3}) ✓ pH = –log[H⁺] = 4.12 ✓ **pH = 5.52** As above for 4.12 but with acid/base ratio inverted. Award 2 OR 1 marks as outlined for 4.12 above

Q6.



Question			Answer	Marks	Guidance
	(c)	(i)	$[CH_3COO^-]$ $n(CH_3COONa) = \frac{9.08}{82.0}$ **OR** 0.111 ✓ (Calc: 0.1107317073) $[CH_3COO^-] = \frac{9.08}{82.0}\times\frac{1000}{250}$ = 0.443 (mol dm^{-3}) **OR** $n(CH_3COOH) = 0.800\times\frac{250}{1000}$ = 0.200 (mol) ✓ **[H⁺]** $[H^+] = K_a\times\frac{[CH_3COOH]}{[CH_3COO^-]}$ **OR** $K_a\times\frac{n(CH_3COOH)}{n(CH_3COO^-)}$ $= 1.75\times10^{-5}\times\frac{0.800}{0.443}$ **OR** $1.75\times10^{-5}\times\frac{0.200}{0.111}$ ✓ $= 3.16\times10^{-5}$ (mol dm^{-3}) ✓ **pH (must come from *calculated* [H⁺])** pH = –log (3.16×10^{-5}) = 4.50 ✓ -- **LAST 3 marks** are **NOT** available using • K_a square root approach (weak acid pH) • K_w /10^{-14} approach (strong base pH) --	**5**	**ALLOW** 2 sig fig **ALLOW** use of **HA** and **A⁻** Mark by **ECF** -- **Alternative method** (If both methods are attempted, mark the method which produces the higher mark) **[H⁺]** $[H^+] = 10^{-pH} = 10^{-4.50}$ $= 3.16\times10^{-5}$ (mol dm^{-3}) ✓ **$[CH_3COO^-]$** $[CH_3COO^-] = K_a\times\frac{[CH_3COOH]}{[H^+]}$ **OR** $1.75\times10^{-5}\times\frac{0.800}{3.16\times10^{-5}}$ ✓ = 0.443 (mol dm^{-3}) ✓ **mass of CH_3COONa** mass $CH_3COONa = 0.443\times\frac{250}{1000}$

			Henderson–Hasselbalch (HH) alternative $pK_a = -\log 1.75 \times 10^{-5} = 4.757$ (or 4.756961951..) $pH = pK_a + \log\frac{[CH_3COO^-]}{[CH_3COOH]}$ **OR** $= pK_a - \log\frac{[CH_3COOH]}{[CH_3COO^-]}$ **OR** $pK_a + \log\frac{0.443}{0.800}$ **OR** $= pK_a - \log\frac{0.800}{0.443}$ ✓ $= pK_a - 0.257$ ✓ $= 4.757 - 0.257 = 4.50$ ✓		**OR** 0.111 ✓ 0.111 × 82.0 = **9.08** (g) ✓ -- **Common errors** **4.64** Use of $M(CH_3COONa) = 60$ 4 marks **2.40** Use of K_a of FCH_2COOH 4 marks