3. ACIDS AND BASES

3.1. pH and $K_w$
3.2. pH and acids
3.3. pH and bases
3.4. Acid-base titrations
3.5. Buffer solutions
3.6. More complex buffer calculations

Acids and bases answers
3.1. pH and $K_w$

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

<table>
<thead>
<tr>
<th>Number ($n$)</th>
<th>$\log_{10} n$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.001</td>
<td>............</td>
</tr>
<tr>
<td>0.1</td>
<td>............</td>
</tr>
<tr>
<td>............</td>
<td>0</td>
</tr>
<tr>
<td>............</td>
<td>3</td>
</tr>
</tbody>
</table>

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

(a) $[H^+] = 1.00 \times 10^{-10}$ mol dm$^{-3}$  
(b) $[H^+] = 0.200 \times 10^{-2}$ mol dm$^{-3}$  
(c) $[H^+] = 3.50 \times 10^{-3}$ mol dm$^{-3}$

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

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

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) $10 \, ^\circ C$, $K_w = 0.29 \times 10^{-14}$ mol$^2$ dm$^{-6}$ ................................................................................................................................................

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

(iii) $40 \, ^\circ C$, $K_w = 2.92 \times 10^{-14}$ 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)

Acids and bases 3.1.
3.2. pH and acids

1. Identify the species formed when the following act as acids;
   (a) HCl ........................................
   (b) NH₄⁺ ....................................
   (c) HCO₃⁻ ....................................

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 NaHSO₄, \( K_a \) of HSO₄⁻ = 1.00 × 10⁻² mol dm⁻³

3. Calculate the concentration of the following acids given their pH.
   (a) HCl, pH 0.65 .................................................................(1 mark)
   (b) H₂SO₄, pH 2.61 .................................................................(1 mark)
   (c) CH₃COOH, pH 3.40, \( K_a \) 1.7 × 10⁻⁵ mol dm⁻³
3.3. 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) $2 \text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4$
   (b) $\text{Ca(OH)}_2 + \text{H}_2\text{CO}_3 \rightarrow \text{CaCO}_3 + 2 \text{H}_2\text{O}$
   (c) $\text{Na}_2\text{CO}_3 + 2 \text{HCl} \rightarrow 2 \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$ .......................................................... (3 marks)

3. Calculate the pH (to 2 dp) of the following basic solutions (take $K_w$ to be $1.00 \times 10^{-14}$ 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 $1.00 \times 10^{-14}$ 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.
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.

<table>
<thead>
<tr>
<th>Titration</th>
<th>Base</th>
<th>Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.100 mol dm$^{-3}$ NaOH</td>
<td>HCl</td>
</tr>
<tr>
<td>2</td>
<td>0.100 mol dm$^{-3}$ KOH</td>
<td>HCOOH</td>
</tr>
<tr>
<td>3</td>
<td>0.100 mol dm$^{-3}$ NH$_3$ solution</td>
<td>HNO$_3$</td>
</tr>
<tr>
<td>4</td>
<td>0.100 mol dm$^{-3}$ NaHCO$_3$</td>
<td>CH$_3$COOH</td>
</tr>
</tbody>
</table>

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): .....................................

Acids and bases 3.4.
3.5. 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$_2$CO$_3$) 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$_2$CO$_3$ is $4.5 \times 10^{-7}$ mol dm$^{-3}$).

   ........................................................................
   ........................................................................
   ........................................................................
   ........................................................................
   ..............................................................................(2 marks)

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

   ........................................................................
   ........................................................................
   ........................................................................
   ........................................................................
   ..............................................................................(2 marks)
3.6. 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 $1.35 \times 10^{-5}$ mol dm$^{-3}$

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.
3. Acids and bases answers

3.1. pH and $K_w$

1. | Number ($n$) | $\log_{10} n$ |
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>0.001</td>
<td>-3</td>
</tr>
<tr>
<td>0.1</td>
<td>-1</td>
</tr>
<tr>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>1.000</td>
<td>3</td>
</tr>
</tbody>
</table>

(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 ($\Delta H$ +ive) so increasing the temperature will favour the forward reaction and hence the [H$^+$] will increase. As a result the pH of the water will decrease as the temperature is increased. (1 mark)
   (b) i) 10 $^\circ$C, $K_w = 0.29 \times 10^{-14}$ mol$^2$ dm$^{-6}$; [H$^+$] = 5.39 $\times$ 10$^{-8}$ mol dm$^{-3}$ $\therefore$ pH = 7.27
   ii) 25 $^\circ$C, $K_w = 1.01 \times 10^{-14}$ mol$^2$ dm$^{-6}$; [H$^+$] = 1.00 $\times$ 10$^{-7}$ mol dm$^{-3}$ $\therefore$ pH = 7.00
   iii) 40 $^\circ$C, $K_w = 2.92 \times 10^{-14}$ mol$^2$ dm$^{-6}$; [H$^+$] = 1.71 $\times$ 10$^{-7}$ mol dm$^{-3}$ $\therefore$ pH = 6.77 (3 marks)
   (c) As the temperature decreases, water remains neutral (1 mark)

   Water is always neutral as [H$^+$] = [OH$^-$] and so there is always an equal number of H$^+$ ions and OH$^-$ ions. (1 mark)

3.2. pH and acids

1. (a) HCl $\rightarrow$ H$^+$ + Cl$^-$
   (b) NH$_4^+$ $\rightleftharpoons$ H$^+$ + NH$_3$
   (c) HCO$_3^-$ $\rightleftharpoons$ H$^+$ + CO$_3^{2-}$ (3 marks)

2. (a) pH = $-\log[0.25]$ = 0.60 (1 mark)
   (b) $K_a = [H^+][SO_3^{2-}] / [HSO_3^-] = [H^+]^2 / [HSO_3^-]$ $\therefore$ $[H^+]^2 = (1.0 \times 10^{-2}) \times 0.004 = 4 \times 10^{-5}$
   $[H^+] = 6.32 \times 10^{-3}$ mol dm$^{-3}$ $\therefore$ pH = $-\log[6.32 \times 10^{-3}]$
   $\therefore$ pH = 2.20 (1 mark for $K_a$ expression, 1 mark for pH)

3. (a) [H$^+$] = 10$^{-0.65}$ $\therefore$ [H$^+$] = 0.22 mol dm$^{-3}$ $\therefore$ [HCl] = 0.22 mol dm$^{-3}$ (1 mark)
   (b) [H$^+$] = 10$^{-2.61}$ $\therefore$ [H$^+$] = 2.45 $\times$ 10$^{-3}$ mol dm$^{-3}$
3. Acids and bases answers

H₂SO₄ → 2 H⁺ + SO₄²⁻ and ∴ [H₂SO₄] = [H⁺]/2 = 1.23 × 10⁻³ mol dm⁻³

(c) CH₃COOH ⇌ H⁺ + CH₃COO⁻ ∴ Kₐ = [H⁺][CH₃COO⁻] = [H⁺]²/[CH₃COOH]
   pH = –log[H⁺], ∴ [H⁺] = 10⁻³.₄⁰, ∴ [H⁺] = 3.98 × 10⁻⁴ mol dm⁻³
   ∴ 1.7 × 10⁻⁵ = [3.98 × 10⁻⁴]² / [CH₃COOH]
   ∴ [CH₃COOH] = 9.32 × 10⁻³ mol dm⁻³

3.3. 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 NH₃ + H₂SO₄ → (NH₄)₂SO₄
   Basic species = NH₃
   (1 mark)
   (b) Ca(OH)₂ + H₂CO₃ → CaCO₃ + 2 H₂O
   Basic species = OH⁻
   (1 mark)
   (c) Na₂CO₃ + 2 HCl → 2 NaCl + H₂O + CO₂
   Basic species = CO₃²⁻
   (3 marks)

3. (a) [NaOH] = 0.150 mol dm⁻³ and ∴ [OH⁻] = 0.150 mol dm⁻³
   ∴ 1 × 10⁻¹⁴ = [H⁺][0.150] and so, [H⁺] = 6.67 × 10⁻¹⁴ mol dm⁻³
   ∴ pH = –log[6.67 × 10⁻¹⁴] = 13.18
   (1 mark)
   (b) [Mg(OH)₂] = 0.261 mol dm⁻³ and ∴ [OH⁻] = 0.261 × 2 = 0.522 mol dm⁻³
   ∴ 1 × 10⁻¹⁴ = [H⁺][0.522] and so, [H⁺] = 1.92 × 10⁻¹⁴ mol dm⁻³
   ∴ pH = –log[1.92 × 10⁻¹⁴] = 13.72
   (1 mark)

4. (a) 11.00 = –log[H⁺] and ∴ [H⁺] = 1.00 × 10⁻¹¹ mol dm⁻³
   ∴ 1.00 × 10⁻¹⁴ = [1.00 × 10⁻¹¹][OH⁻] and so, [OH⁻] = 1.00 × 10⁻³ mol dm⁻³
   ∴ [KOH] = 1.00 × 10⁻³ mol dm⁻³
   (1 mark)
   (b) 10.45 = –log[H⁺] and ∴ [H⁺] = 3.55 × 10⁻¹¹ mol dm⁻³
   ∴ 1.00 × 10⁻¹⁴ = [3.55 × 10⁻¹¹][OH⁻] and so, [OH⁻] = 2.82 × 10⁻⁴ mol dm⁻³
   Since 1 mol of Ca(OH)₂ contains 2 mol OH⁻, [Ca(OH)₂] = 1.41 × 10⁻⁴ mol dm⁻³
   (1 mark)

5. [Diagram of chemical reaction]
3. Acids and bases answers

3.4. Acid-base titrations

Titration number: 1 2 3 4
Suitable indicator: phenolphthalein or methyl orange
Conc. of acid (if needed):
\[
\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}
\]
Moles in 25 \(\text{cm}^3\) 0.100 mol dm\(^{-3}\) \(\text{NaOH} = 2.5 \times 10^{-3}\)
Volume of HCl needed for neutralisation = 14 cm\(^3\)
Conc. of HCl = \(2.5 \times 10^{-3} / 0.014 \text{ dm}^3 = 0.18 \text{ mol dm}^{-3}\)

(1 mark for correct identification of each titration, 1 mark for each suitable indicator named, 1 mark for each calculation of acid concentration)
3. Acids and bases answers

3.5. Buffer solutions

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

$\text{pK}_a = -\log K_a, \therefore K_a = 10^{-3.75} = 1.78 \times 10^{-4} \text{ mol dm}^{-3}$  

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

$[\text{HCOOH}(aq)] = 0.450 \text{ mol} / 0.5 \text{ dm}^3 = 0.90 \text{ mol dm}^{-3}$ 

$[\text{HCOOH}(aq)] = 0.510 \text{ mol} / 0.5 \text{ dm}^3 = 1.02 \text{ mol dm}^{-3}$ 

Substituting these values in we get, $1.78 \times 10^{-4} \text{ mol dm}^{-3} = 0.90 \times [\text{H}^+(aq)] / 1.02$ 

$\therefore [\text{H}^+(aq)] = 2.02 \times 10^{-4} \text{ mol dm}^{-3}$ 

$\text{pH} = 3.70$  

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

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

$\text{HCOOH} + \text{OH}^- \rightarrow \text{HCOO}^- + \text{H}_2\text{O}$ 

This removes the $\text{OH}^-$ and so the pH remains approximately constant.

2. (a) $\text{H}_2\text{CO}_3(aq) \rightleftharpoons \text{HCO}_2^- (aq) + \text{H}^+(aq)$ 

$p\text{H}^+ = 7.40$, so $[\text{H}^+(aq)] = 10^{-7.40} = 3.98 \times 10^{-8} \text{ mol dm}^{-3}$  

$K_a = \frac{[\text{HCO}_2^- (aq)][\text{H}^+(aq)]}{[\text{H}_2\text{CO}_3(aq)]}$ 

$[\text{H}_2\text{CO}_3(aq)] = 3.98 \times 10^{-8} \text{ mol dm}^{-3}$ 

Since both stock solutions are of an equal concentration they should mix the two in a ratio of $11.3 : 1 \text{ HCO}_2^- : \text{H}_2\text{CO}_3$.

(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.

3.6. More complex buffer calculations

1. $\text{CH}_3\text{CH}_2\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{CH}_2\text{COO}^-\text{Na}^+ + \text{H}_2\text{O}$ 

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

$\therefore$ moles of $\text{CH}_3\text{CH}_2\text{COOH}$ will decrease by $1.5 \times 10^{-3} \text{ mol}$ and moles of $\text{CH}_3\text{CH}_2\text{COO}^-\text{Na}^+$ will increase by $1.5 \times 10^{-3} \text{ mol}$.
3. Acids and bases answers

\[
\text{CH}_3\text{CH}_2\text{COOH} \rightleftharpoons \text{CH}_3\text{CH}_2\text{COO}^- + \text{H}^+
\]

<table>
<thead>
<tr>
<th>Initial moles</th>
<th>Change in moles</th>
<th>Equilibrium moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>(0.035 \text{ dm}^3 \times 0.150 \text{ mol dm}^{-3})</td>
<td>(-1.5 \times 10^{-3} \text{ mol})</td>
<td>(3.75 \times 10^{-3} \text{ mol})</td>
</tr>
<tr>
<td>0 mol</td>
<td>1.5 \times 10^{-3} \text{ mol}</td>
<td>?</td>
</tr>
<tr>
<td>0 mol</td>
<td>?</td>
<td>?</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[\text{CH}_3\text{CH}_2\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{CH}_2\text{COOH}]} = \frac{(1.5 \times 10^{-3} \text{ mol} / 0.05 \text{ dm}^3) \times [\text{H}^+]}{(3.75 \times 10^{-3} \text{ mol} / 0.05 \text{ dm}^3)} = 1.35 \times 10^{-5} \text{ mol dm}^{-3}
\]

\[
[\text{H}^+] = \frac{3.75 \times 10^{-3} \text{ mol}}{0.05 \text{ dm}^3} = 7.5 \times 10^{-5} \text{ mol dm}^{-3}
\]

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

\[
\text{pH} = 4.42
\]

2. In a 10 cm\(^3\) aliquot (\(= 1/5 \text{ th}\)) of the buffer solution made above;

moles of \(\text{CH}_3\text{CH}_2\text{COOH}\) = 7.5 \(\times 10^{-4}\) mol; moles of \(\text{CH}_3\text{CH}_2\text{COO}^-\) = 3.0 \(\times 10^{-4}\) mol

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

\[
\text{CH}_3\text{CH}_2\text{COO}^- + \text{H}^+ \rightarrow \text{CH}_3\text{CH}_2\text{COOH}
\]

\[
\text{CH}_3\text{CH}_2\text{COOH} \rightleftharpoons \text{CH}_3\text{CH}_2\text{COO}^- + \text{H}^+
\]

<table>
<thead>
<tr>
<th>Initial moles</th>
<th>Change in moles</th>
<th>Equilibrium moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>(7.5 \times 10^{-4}) mol</td>
<td>(+2.5 \times 10^{-5}) mol</td>
<td>(7.75 \times 10^{-4}) mol</td>
</tr>
<tr>
<td>(3.0 \times 10^{-4}) mol</td>
<td>(-2.5 \times 10^{-5}) mol</td>
<td>(2.75 \times 10^{-4}) mol</td>
</tr>
</tbody>
</table>

\[
[\text{H}^+] = \frac{(2.75 \times 10^{-4} \text{ mol} / 0.0105 \text{ dm}^3) \times [\text{H}^+]}{(7.75 \times 10^{-4} \text{ mol} / 0.0105 \text{ dm}^3)} = 3.80 \times 10^{-5} \text{ mol dm}^{-3}
\]

\[
\text{pH} = 4.42
\]

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

\[
\text{CH}_3\text{CH}_2\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{CH}_2\text{COO}^- + \text{H}_2\text{O}
\]

\[
\text{no. of moles of OH}^- \text{ added} = 2 \times 2.5 \times 10^{-5} \text{ mol} = 5.0 \times 10^{-5} \text{ mol}
\]

\[
\text{moles of CH}_3\text{CH}_2\text{COOH} \text{ will decrease by } 5.0 \times 10^{-5} \text{ mol and moles of CH}_3\text{CH}_2\text{COO}^- \text{ will increase by } 5.0 \times 10^{-5} \text{ mol.}
\]
3. Acids and bases answers

\[ \text{CH}_3\text{CH}_2\text{COOH} \rightleftharpoons \text{CH}_3\text{CH}_2\text{COO}^- + \text{H}^+ \]

**Initial moles**
- \[7.5 \times 10^{-4} \text{ mol}\]
- \[3.0 \times 10^{-4} \text{ mol}\]

**Change in moles**
- \[-5.0 \times 10^{-5} \text{ mol}\]
- \[+5.0 \times 10^{-5} \text{ mol}\]

**Equilibrium moles**
- \[7.0 \times 10^{-4} \text{ mol}\]
- \[3.5 \times 10^{-4} \text{ mol}\]

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

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

\[ \therefore \text{pH} = 4.57 \]

(1 mark)