Acids and bases
pH and $K_w$

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

<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: (3 marks)

(a) $[H^+] = 1.00 \times 10^{-10}$ mol dm$^{-3}$, pH = ........

(b) $[H^+] = 0.200 \times 10^{-2}$ mol dm$^{-3}$, pH = ........

(c) $[H^+] = 3.50 \times 10^{-3}$ mol dm$^{-3}$, pH = ........

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

$$\text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq) \quad \Delta H = +\text{ive}$$

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, $K_w$;

$$K_w = [\text{H}^+(aq)] [\text{OH}^-(aq)]$$

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$^{-6}$

(c) Complete the paragraph below;

As the temperature decreases, water becomes (more acidic / less acidic / remains neutral). Explain your answer (2 marks)
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, Kₐ of HSO₄⁻ = 1.00 × 10⁻² mol dm⁻³ ________________________________ (2 marks)

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, Kₐ 1.7 × 10⁻⁵ mol dm⁻³ ________________________________ (2 marks)
**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) \(2 \text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4^+)\text{SO}_4^{2-}\)
   (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.

![Diagram of ethylamine reacting with a proton](image-url)
Acid-base tritrations

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>

Titrations: 1 2 3 4

Suitable indicator: ......................................

Conc. of acid (if needed): ..................................

Titrations: 1 2 3 4

Suitable indicator: ......................................

Conc. of acid (if needed): ..................................
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)
**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$^{-3}$)

(4 marks)

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.

(6 marks)
Acids and bases – Answers

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>1000</td>
<td>$3$</td>
</tr>
</tbody>
</table>

   (1 mark for all numbers correct)

2. (a) $\text{pH} = 10.00$
    (b) $\text{pH} = 2.70$
    (c) $\text{pH} = 2.46$

   (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 $[H^+(aq)]$ will increase. As a result the pH of the water will decrease as the temperature is increased.

   (1 mark)

   (b)
   (i) $10 \degree C$, $K_w = 0.29 \times 10^{-14} \text{ mol dm}^{-6}$; $[H^+(aq)] = 5.39 \times 10^{-8} \text{ mol dm}^{-3} \therefore \text{pH} = 7.27$
   (ii) $25 \degree C$, $K_w = 1.01 \times 10^{-14} \text{ mol dm}^{-6}$; $[H^+(aq)] = 1.00 \times 10^{-7} \text{ mol dm}^{-3} \therefore \text{pH} = 7.00$
   (iii) $40 \degree C$, $K_w = 2.92 \times 10^{-14} \text{ mol dm}^{-6}$; $[H^+(aq)] = 1.71 \times 10^{-7} \text{ mol dm}^{-3} \therefore \text{pH} = 6.77$

   (3 marks)

   (c) As the temperature decreases, water remains neutral

   (1 mark)

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

(1 mark)
pH and acids

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

   (3 marks)

2. (a) \( \text{pH} = -\log[0.25] = 0.60 \)

   (1 mark)

   (b) \( K_a = \frac{[\text{H}^+][\text{SO}_4^{2-}]}{[\text{HSO}_4^-]} \) \( \therefore [\text{H}^+]^2 = (1.0 \times 10^{-2}) \times 0.004 = 4 \times 10^{-5} \)

   \( \therefore [\text{H}^+] = 6.32 \times 10^{-3} \) mol dm\(^{-3} \)

   \( \therefore \text{pH} = -\log[6.32 \times 10^{-3}] \)

   \( \therefore \text{pH} = 2.20 \)

   (1 mark for \( K_a \) expression, 1 mark for pH)

3. (a) \( [\text{H}^+] = 10^{-0.65} \therefore [\text{H}^+] = 0.22 \) mol dm\(^{-3} \) \( \therefore [\text{HCl}] = 0.22 \) mol dm\(^{-3} \)

   (1 mark)

   (b) \( [\text{H}^+] = 10^{-2.61} \therefore [\text{H}^+] = 2.45 \times 10^{-3} \) mol dm\(^{-3} \)

   \( \text{H}_2\text{SO}_4 \rightarrow 2 \text{H}^+ + \text{SO}_4^{2-} \) and \( \therefore [\text{H}_2\text{SO}_4] = [\text{H}^+] / 2 = 1.23 \times 10^{-3} \) mol dm\(^{-3} \)

   (1 mark)

   (c) \( \text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^- \) \( \therefore K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \) \( \therefore [\text{H}^+]^2 = [\text{H}^+] \times [\text{CH}_3\text{COO}^-] / [\text{CH}_3\text{COOH}] \)

   \( \text{pH} = -\log[H^+] \therefore [\text{H}^+] = 10^{-3.40} \therefore [\text{H}^+] = 3.98 \times 10^{-4} \) mol dm\(^{-3} \)

   (1 mark)

   \( \therefore 1.7 \times 10^{-5} = [3.98 \times 10^{-4}]^2 / [\text{CH}_3\text{COOH}] \)

   \( \therefore [\text{CH}_3\text{COOH}] = 9.32 \times 10^{-3} \) mol dm\(^{-3} \)

   (1 mark)
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) $2 \text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)^+\text{SO}_4^{2-}$  Basic species = $\text{NH}_3$
   (1 mark)

   (b) $\text{Ca(OH)}_2 + \text{H}_2\text{CO}_3 \rightarrow \text{CaCO}_3 + 2 \text{H}_2\text{O}$  Basic species = $\text{OH}^-$
   (1 mark)

   (c) $\text{Na}_2\text{CO}_3 + 2 \text{HCl} \rightarrow 2 \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$  Basic species = $\text{CO}_3^{2-}$
   (3 marks)

3. (a) $[\text{NaOH}] = 0.150 \text{ mol dm}^{-3}$ and ∴ $[\text{OH}^-] = 0.150 \text{ mol dm}^{-3}$

   ∴ $1 \times 10^{-14} = [\text{H}^+][0.150]$ and so, $[\text{H}^+] = 6.67 \times 10^{-14} \text{ mol dm}^{-3}$

   ∴ pH = $-\log[6.67 \times 10^{-14}] = 13.18$
   (1 mark)

   (b) $[\text{Mg(OH)}_2] = 0.261 \text{ mol dm}^{-3}$ and ∴ $[\text{OH}^-] = 0.261 \times 2 = 0.522 \text{ mol dm}^{-3}$

   ∴ $1 \times 10^{-14} = [\text{H}^+][0.522]$ and so, $[\text{H}^+] = 1.92 \times 10^{-14} \text{ mol dm}^{-3}$

   ∴ pH = $-\log[1.92 \times 10^{-14}] = 13.72$
   (1 mark)

4. (a) $11.00 = -\log[\text{H}^+]$ and ∴ $[\text{H}^+] = 1.00 \times 10^{-11} \text{ mol dm}^{-3}$

   ∴ $1.00 \times 10^{-14} = [1.00 \times 10^{-11}][\text{OH}^-]$ and so, $[\text{OH}^-] = 1.00 \times 10^{-3} \text{ mol dm}^{-3}$

   ∴ $[\text{KOH}] = 1.00 \times 10^{-3} \text{ mol dm}^{-3}$
   (1 mark)

   (b) $10.45 = -\log[\text{H}^+]$ and ∴ $[\text{H}^+] = 3.55 \times 10^{-11} \text{ mol dm}^{-3}$

   ∴ $1.00 \times 10^{-14} = [3.55 \times 10^{-11}][\text{OH}^-]$ and so, $[\text{OH}^-] = 2.82 \times 10^{-4} \text{ mol dm}^{-3}$

   Since 1 mol of $\text{Ca(OH)}_2$ contains 2 mol $\text{OH}^-$, $[\text{Ca(OH)}_2] = 1.41 \times 10^{-4} \text{ mol dm}^{-3}$
   (1 mark)

5. 

   (1 mark)
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 cm\(^3\) 0.100 mol dm\(^{-3}\) 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 dm\(^3\) = 0.18 mol dm\(^{-3}\)

Titration number: 1 2 3 4

Suitable indicator: methyl orange

Conc. of acid (if needed): N/A

Titration number: 1 2 3 4

Suitable indicator: phenolphthalein

Conc. of acid (if needed):
\[
\text{KOH} + \text{HCOOH} \rightarrow \text{HCOO}^-\text{K}^+ + \text{H}_2\text{O}
\]
Moles in 25 cm\(^3\) 0.100 mol dm\(^{-3}\) KOH = 2.5 \times 10^{-3}
Volume of HCOOH needed for neutralisation = 42 cm\(^3\)
Conc. of HCOOH = 2.5 \times 10^{-3} / 0.042 dm\(^3\) = 0.060 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)
**Buffer solutions**

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

(ii) \( pK_a = -\log K_a, \therefore K_a = 10^{-3.75} = 1.78 \times 10^{-4} \text{ mol dm}^{-3} \)  
   \[ (1 \text{ mark}) \]

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

\[ [\text{HCOO}^-(aq)] = 0.450 \text{ mol / 0.5 dm}^3 = 0.90 \text{ mol dm}^{-3} \]
\[ [\text{HCOOH(aq)}] = 0.510 \text{ mol / 0.5 dm}^3 = 1.02 \text{ mol dm}^{-3} \]

Substituting these values in we get, \( 1.78 \times 10^{-4} \text{ mol dm}^{-3} = 0.90 \times [\text{H}^+(aq)] / 1.02 \)
\[ \therefore [\text{H}^+(aq)] = 2.02 \times 10^{-4} \text{ mol dm}^{-3} \]  
   \[ (1 \text{ mark}) \]

\[ \therefore \text{pH} = 3.70 \]  
   \[ (1 \text{ mark}) \]

(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.  
   \[ (1 \text{ mark}) \]

(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.  
   \[ (1 \text{ mark}) \]

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

pH of desired buffer = 7.40, so \([\text{H}^+(aq)] = 10^{-7.40} = 3.98 \times 10^{-8} \text{ mol dm}^{-3} \)  
   \[ (1 \text{ mark}) \]

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

\[ \therefore [\text{HCO}_3^-(aq)] = K_a = 4.5 \times 10^{-7} \text{ mol dm}^{-3} = 11.3 \]
\[ \frac{[\text{H}_2\text{CO}_3(aq)]}{[\text{H}^+(aq)]} = \frac{3.98 \times 10^{-8} \text{ mol dm}^{-3}}{1} \]

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.  
   \[ (2 \text{ marks}) \]
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 dm\(^3\) \times 0.100 mol dm\(^{-3}\) = 1.5 \times 10^{-3} mol

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

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

\[
\therefore [\text{H}^+] = 3.38 \times 10^{-5} \text{ mol dm}^{-3}
\]

\[
\therefore \text{pH} = 4.47
\]

2. In a 10 cm\(^3\) aliquot (= 1/5 th) of the buffer solution made above;
moles of CH\(_3\)CH\(_2\)COOH = 7.5 \times 10^{-4} \text{ mol}; moles of CH\(_3\)CH\(_2\)COO\(^-\) = 3.0 \times 10^{-4} \text{ mol}

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

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

\[
\therefore \text{moles of CH}_3\text{CH}_2\text{COOH will increase by } 2.5 \times 10^{-5} \text{ mol and moles of CH}_3\text{CH}_2\text{COO}^- \text{ will decrease by } 2.5 \times 10^{-5} \text{ mol.}
\]

\[
K_a = [\text{CH}_3\text{CH}_2\text{COO}^-][\text{H}^+] = \frac{(2.75 \times 10^{-4} \text{ mol} / 0.0105 \text{ dm}^3) \times [\text{H}^+]}{(7.75 \times 10^{-4} \text{ mol} / 0.0105 \text{ dm}^3)}
\]

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

\[
\therefore \text{pH} = 4.42
\]
(b) No. of moles of Ca(OH)$_2$ added = 0.0005 dm$^3 \times 0.05$ mol dm$^{-3} = 2.5 \times 10^{-5}$ mol  
\[\Rightarrow\] no. of moles of OH$^-$ added = 2 $\times$ 2.5 $\times$ 10$^{-5}$ mol = 5.0 $\times$ 10$^{-5}$ mol  

\[\text{CH}_3\text{CH}_2\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{CH}_2\text{COO}^- + \text{H}_2\text{O}\]
\[\Rightarrow\] moles of CH$_3$CH$_2$COOH will decrease by 5.0 $\times$ 10$^{-5}$ mol and moles of CH$_3$CH$_2$COO$^-$ will increase by 5.0 $\times$ 10$^{-5}$ mol.

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

<table>
<thead>
<tr>
<th>Initial moles</th>
<th>7.5 $\times$ 10$^{-4}$ mol</th>
<th>3.0 $\times$ 10$^{-4}$ mol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change in moles</td>
<td>$-5.0 \times 10^{-5}$ mol</td>
<td>$+5.0 \times 10^{-5}$ mol</td>
</tr>
<tr>
<td>Equilibrium moles</td>
<td>7.0 $\times$ 10$^{-4}$ mol</td>
<td>3.5 $\times$ 10$^{-4}$ mol</td>
</tr>
</tbody>
</table>

\[1.35 \times 10^{-5} \text{ mol dm}^{-3} = \frac{(3.5 \times 10^{-4} \text{ mol / 0.0105 dm}^3) \times [\text{H}^+]}{(7.0 \times 10^{-4} \text{ mol / 0.0105 dm}^3)}\]
\[\Rightarrow [\text{H}^+] = 2.7 \times 10^{-5} \text{ mol dm}^{-3}\]
\[\Rightarrow \text{pH} = 4.57\]