Raffles Institution
Year 6 H2 Chemistry 2026
Tutorial 15 – Acid-Base Equilibria

Suggested Answers to Practice Questions

1 (a) $\mathbf{pH = -\lg [H^+]}$

$[H^+] = 10^{-3.50} =$ **$\underline{\mathbf{3.16 \times 10^{-4}\ mol\ dm^{-3}}}$**

(b) $HA(aq) + NaOH(aq) \rightarrow NaA(aq) + H_2O(l)$

$$n_{NaOH}\text{ reacted} = n_{HA}\text{ present in apple juice} = \frac{27.50}{1000} \times 0.100 = 2.75 \times 10^{-3}\text{ mol}$$

$$[HA]\text{ in apple juice} = \frac{2.75 \times 10^{-3}}{25.0} \times 1000 = 0.110\text{ mol dm}^{-3}$$

Since **$[H^+] << [HA]$**, the acid HA in apple juice undergoes only **partial dissociation** and is therefore a **weak acid**.

(c) $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$

$$\text{degree of dissociation of HA} = \frac{n_{HA}\text{ dissociated in 1 dm}^3\text{ of apple juice}}{\text{initial } n_{HA}\text{ in 1 dm}^3\text{ of apple juice}}$$

$$= \frac{3.16 \times 10^{-4}}{0.110} = \underline{\mathbf{2.87 \times 10^{-3}}}$$

At equilibrium, $[H_3O^+] = [A^-] = 3.16 \times 10^{-4}$ mol dm^{-3}
$[HA] = 0.110 - 3.16 \times 10^{-4} = 0.1097$ mol dm^{-3}

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(3.16 \times 10^{-4})^2}{0.1097} = \underline{\mathbf{9.10 \times 10^{-7}\ mol\ dm^{-3}}}$$

(d) Since both acids have the same pH (3.50), both acids have the same $[H^+]$. However, partial dissociation occurs in apple juice, hence the total acid concentration is higher (than hydrochloric acid).
$[HCl] = [H^+] = 3.16 \times 10^{-4}$ mol dm^{-3}
From (b): [HA] in apple juice = 0.110 mol dm^{-3}
In 1 dm^3, the sample of apple juice will have a greater amount of acid present. Since the volume of H_2 to be produced is proportional to the amount of acid present and a **greater amount of acid** is present in the sample of apple juice, the **apple juice** sample would yield a larger volume of hydrogen gas than the hydrochloric acid sample.

2 (a) (i) $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

At equilibrium, $[H^+] = [CH_3COO^-]$
$[CH_3COOH]_{eqm} \approx [CH_3COOH]_{initial} = 0.50$ mol dm^{-3},
since CH_3COOH is a weak acid with a small K_a

$$K_a = \frac{\left[CH_3COO^-\right]\left[H^+\right]}{\left[CH_3COOH\right]} = \frac{\left[H^+\right]^2}{\left[CH_3COOH\right]}$$

$$\left[H^+\right] = \sqrt{K_a \times \left[CH_3COOH\right]} = \sqrt{10^{-4.74} \times 0.50} = 3.016 \times 10^{-3} \text{ mol dm}^{-3}$$

$$pH = -\lg\left(3.016 \times 10^{-3}\right) = \underline{\mathbf{2.52}}$$

(ii) Let volume of CH_3COOH used be V dm^3.

Initial $n_{CH_3COOH} = 0.50V$ mol Initial $n_{NaOH} = 0.20V$ mol

Upon mixing, the following reaction takes place:

$CH_3COOH + NaOH \rightarrow CH_3COO^-Na^+ + H_2O$

Hence CH_3COOH is in excess.

$n_{CH_3COO^-}$ formed $= n_{NaOH}$ reacted $= 0.20V$ mol

Unreacted $n_{CH_3COOH} = 0.50V - 0.20V = 0.30V$ mol

$\Rightarrow$ The resultant solution is a buffer.

$$pH \text{ of resultant solution} = pK_a + \lg\frac{\left[CH_3COO^-\right]}{\left[CH_3COOH\right]} = 4.74 + \lg\frac{\left(\frac{0.20V}{2V}\right)}{\left(\frac{0.30V}{2V}\right)} = \underline{\mathbf{4.56}}$$

(b) (i) $CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$

CH_3COO^- undergoes <u>hydrolysis in water to form OH^-</u>, which causes the solution to be alkaline and pH is above 7. (Note: $[OH^-] > [H_3O^+]$)

(ii) $$K_b = \frac{\left[CH_3COOH\right]\left[OH^-\right]}{\left[CH_3COO^-\right]}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{10^{-4.74}} = \underline{5.50\times10^{-10} \text{ mol dm}^{-3}}$$

(iii) $CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$

At equilibrium, $[OH^-] = [CH_3COOH]$

$[CH_3COO^-]_{eqm} \approx [CH_3COO^-]_{initial} = 0.050$ mol dm^{-3}, since CH_3COO^- is a weak base with a small K_b

$$K_b = \frac{\left[CH_3COOH\right]\left[OH^-\right]}{\left[CH_3COO^-\right]} = \frac{\left[OH^-\right]^2}{\left[CH_3COO^-\right]}$$

$$\left[OH^-\right] = \sqrt{K_b \times \left[CH_3COO^-\right]} = \sqrt{5.50\times10^{-10} \times 0.050} = 5.242\times10^{-6} \text{ mol dm}^{-3}$$

$$pOH = -\lg\left(5.242\times10^{-6}\right) = 5.28$$

$$pH = 14 - 5.28 = \underline{\mathbf{8.72}}$$

(c) Let the volume of $CH_3COO^-Na^+(aq)$ needed be V dm^3.
Amount of CH_3COO^- = V x 0.10 = 0.10 V mol.
Amount of CH_3COOH = 0.025 x 0.12 = 0.0030 mol

$$\text{pH of resultant solution} = pK_a + \lg\frac{[CH_3COO^-]}{[CH_3COOH]}$$

$$5.00 = 4.74 + \lg\frac{\left(\frac{0.10V}{V_{total}}\right)}{\left(\frac{0.0030}{V_{total}}\right)}$$

$$\lg\left(\frac{0.10V}{0.0030}\right) = 0.26$$

$$V = 10^{0.26} \times \frac{0.0030}{0.10} = \underline{\mathbf{0.0546\ dm^3}} = \underline{\mathbf{54.6\ cm^3}}$$

3 **(a)** The resultant solution is a buffer.

$$n_{NH_4^+}\ \text{present} = \frac{750}{1000} \times 0.20 = 0.15\ \text{mol}$$

$$n_{NH_3}\ \text{present} = \frac{500}{1000} \times 0.10 = 0.05\ \text{mol}$$

At equilibrium in the resultant solution, $[NH_4^+]_{eqm} \approx [NH_4^+]_{initial}$ and $[NH_3]_{eqm} \approx [NH_3]_{initial}$

$$pK_b = -\lg(\frac{K_w}{K_a}) = -\lg(\frac{1.0 \times 10^{-14}}{6.00 \times 10^{-10}}) = 4.778$$

$$\text{pOH of resultant solution} = pK_b + \lg\frac{[NH_4^+]}{[NH_3]} = 4.778 + \lg\frac{\left(\frac{0.15}{V_{total}}\right)}{\left(\frac{0.05}{V_{total}}\right)} = 4.778 + \lg\frac{0.15}{0.05} = 5.255$$

pH = 14 – 5.255 = 8.745 = **<u>8.75</u>**

(b) (i) Upon addition of OH^-, the following reaction occurs:
$NH_4^+(aq)$ + $OH^-(aq)$ → $NH_3(aq)$ + $H_2O(l)$

$$n_{NH_4^+}\ \text{present in resultant solution} = 0.15 - 0.002 = 0.148\ \text{mol}$$

$$n_{NH_3}\ \text{present in resultant solution} = 0.05 + 0.002 = 0.052\ \text{mol}$$

The resultant solution is still a buffer.

$$\text{pOH of resultant solution} = pK_b + \lg\frac{[NH_4^+]}{[NH_3]} = 4.778 + \lg\frac{\left(\frac{0.148}{V_{total}}\right)}{\left(\frac{0.052}{V_{total}}\right)} = 4.778 + \lg\frac{0.148}{0.052} = 5.232$$

pH = 14 – 5.232 = 8.768

Change in pH = 8.768 – 8.745 = **<u>+0.023</u>**
⇒ There is an **increase in pH of 0.023 units**.

(ii) n_{H^+} added $= \frac{1.0}{1000} \times 2.00 = 0.002$ mol

Upon addition of H^+, the following reaction occurs:
$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$

$n_{NH_4^+}$ present in resultant solution $= 0.15 + 0.002 = 0.152$ mol

n_{NH_3} present in resultant solution $= 0.05 - 0.002 = 0.048$ mol

The resultant solution is still a buffer.

pOH of resultant solution $= pK_b + \lg \frac{[NH_4^+]}{[NH_3]} = 4.778 + \lg \frac{\left(\frac{0.152}{V_{total}}\right)}{\left(\frac{0.048}{V_{total}}\right)} = 4.778 + \lg \frac{0.152}{0.048} = 5.279$

pH = 14 – 5.279 = 8.721

Change in pH = 8.721 – 8.745 = **–0.024**
$\Rightarrow$ There is a **decrease in pH of 0.024 units**.

4 **(a)** at pH = 8.2, $[H^+] = 10^{-8.2} = 6.3096 \times 10^{-9}$
At pH = 8.1, $[H^+] = 10^{-8.1} = 7.9433 \times 10^{-9}$
% change in $[H^+]$
$= (7.9433 - 6.3096)/6.3096 \times 100\%$
= 25.9%

(b) As the amount of CO_2 gas increases in the atmosphere, $[CO_2(aq)]$ increases. This causes the position of equilibrium in reaction 4.1 to shift to the right, increasing $[H^+]$. To counteract the increase in $[H^+]$, position of equilibrium in reaction 4.2 shifts left, partially offsetting the increase in H^+ concentration, thus maintaining the pH in seawater.
In other words, with respect to reaction 4.2, the backward reaction takes place to neutralise additional H^+, minimising the change in $[H^+]$ and resisting pH change.

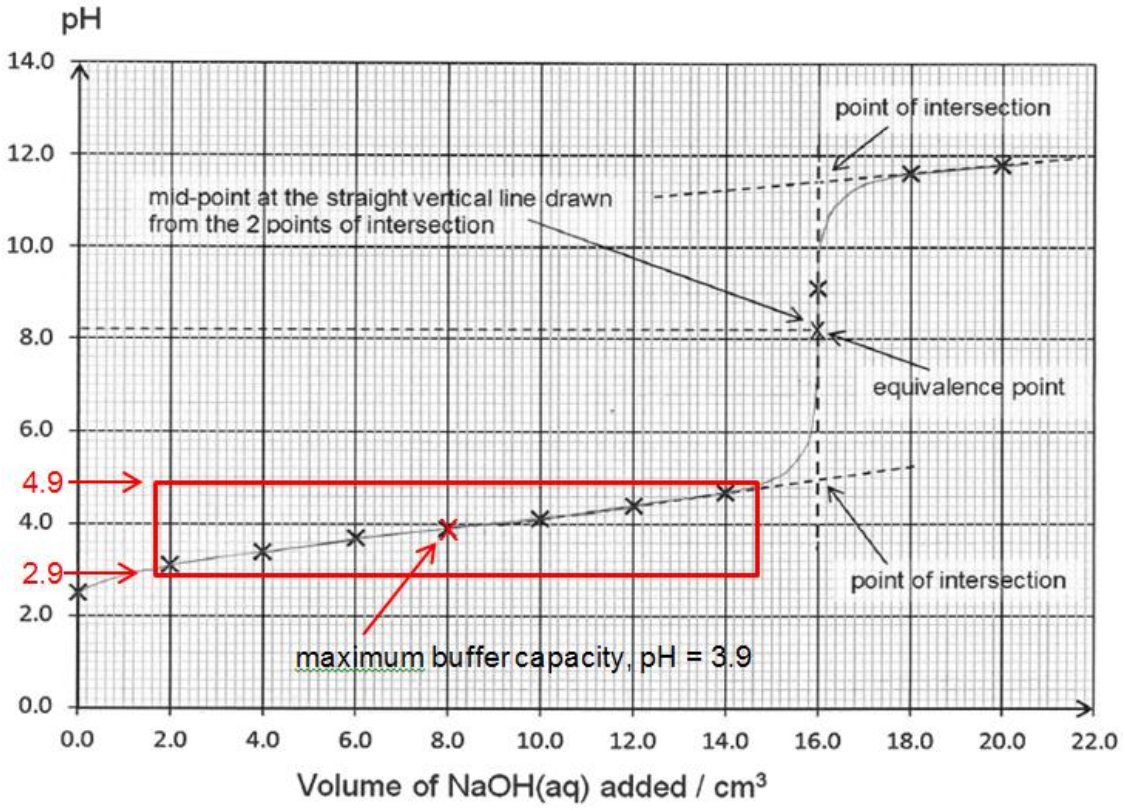
$$CO_3^{2-} + H^+ \longrightarrow HCO_3^-$$

(c) At pH = 8.1, $[H^+] = 10^{-8.1} = 7.9433 \times 10^{-9}$
From reaction 6.1,
$[HCO_3^-]/[CO_2] = K_1/[H^+] = 182.5$
HCO_3^- is more abundant than CO_2
From reaction 6.2,
$[CO_3^{2-}]/[HCO_3^-] = K_2/[H^+] = 0.1385$
HCO_3^- is more abundant than CO_3^{2-}
Hence, HCO_3^- is most abundant inorganic carbon species at pH 8.1.

5 (Ans: D)

Option **A**	Incorrect. The correct pH of the buffer solution is: $pH = pK_a + \lg \frac{[\text{lactate}]}{[\text{lactic acid}]} = -\lg(1.4 \times 10^{-4}) + \lg \frac{0.5}{1.5} = 3.38$
Option **B**	Incorrect. Upon dilution with some water, the pH of the buffer solution should remain the same as the amount of lactic acid and sodium lactate in the buffer solution remains unchanged (although their concentrations decrease). $pH = pK_a + \lg \frac{[\text{lactate}]}{[\text{lactic acid}]} = pK_a + \lg \frac{n(\text{lactate}) \,/\, \cancel{V_{buffer}}}{n(\text{lactic acid}) \,/\, \cancel{V_{buffer}}} = pK_a + \lg \frac{n(\text{lactate})}{n(\text{lactic acid})}$ *(If the buffer solution is infinitely diluted with water, the $[H^+]$ in the buffer solution tends towards that of pure water (i.e. 10^{-7} mol dm^{-3}) and pH increases and approaches 7.)* Hence, pH of the buffer solution will not decrease upon dilution with water.
Option **C**	Incorrect. Buffering capacity depends on the amounts of lactic acid and sodium lactate in the buffer solution. Since the amounts of lactic acid and sodium lactate in the buffer solution remain unchanged upon dilution, the buffering capacity remains the same.
Option **D**	Correct. Since the buffer solution has a greater amount of lactic acid (weak acid) than the lactate ions (conjugate base), the buffer solution is more effective in buffering the effect of addition of small amounts of base than acid.

6 **(a)**



[Note: marking out the equivalence point on the graph is not required]

The two major organic species are $CH_3CH(OH)CO_2H$ and $CH_3CH(OH)CO_2^-$.

In the solution, the following equilibria exist,

$CH_3CH(OH)CO_2H(aq) \rightleftharpoons CH_3CH(OH)CO_2^-(aq) + H^+(aq)$ .....(1)
$CH_3CH(OH)CO_2^-(aq) + H_2O(l) \rightleftharpoons CH_3CH(OH)CO_2H(aq) + OH^-(aq)$ .....(2)

When a small amount of OH^- ions is added to the solution, concentration of OH^- increases. By Le Chatelier's Principle, the position of equilibrium 2 shift to the left to partially offset the increase in concentration of OH^- by favouring the backward reaction of the equilibrium, resulting in a small increase in pH.

(b) (i) **Thymol blue (or thymolphthalein)** may be used for this **weak acid–strong base titration**.

Its **pH range** (around pH 8.0 – 9.6, or 9.4 – 10.6 for thymolphthalein) **coincides** with the region of the **rapid pH change** (about pH 7.5 – 10.5) of the titration. Thus it will **change colour sharply** when near the equivalence point.

(ii) The colour change will be from **yellow to green (or colourless to light blue for thymolphthalein)**.

(c) $CH_3CH(OH)CO_2H(aq) + NaOH(aq) \rightarrow CH_3CH(OH)CO_2^-Na^+(aq) + H_2O(l)$

From the graph, volume of NaOH required to reach equivalence point = **16.00 cm³**

$n_{\text{lactic acid}}$ present $= n_{NaOH}$ used for neutralisation $= \frac{16.00}{1000} \times 0.050 = 8.00 \times 10^{-4}$ mol

$$[\text{lactic acid}] = \frac{8.00 \times 10^{-4}}{10.0} \times 1000 = \underline{\mathbf{0.0800\ mol\ dm^{-3}}}$$

[Note: Accept all sensible answers (e.g. 15.80 cm^3) based on the titration curve drawn. Some titration curves will be drawn without a vertical straight line portion.]

(d) From **(c)**, volume of NaOH needed to neutralise half of the lactic acid present = (½)(16.00) = **8.00 cm^3**

When 8.00 cm^3 of NaOH is added, the resultant solution is a buffer with maximum buffer capacity, i.e. $[CH_3CH(OH)CO_2H] = [CH_3CH(OH)CO_2^-]$ and **pH = pK_a = 3.9**.

Hence $K_a = 10^{-3.9}$ = **<u>1.26 x 10^{-4} mol dm^{-3}</u>**

(e) At the equivalence point of the titration, the resultant solution is **$CH_3CH(OH)CO_2^-Na^+$(aq).**

Being the conjugate base of a weak acid, $CH_3CH(OH)CO_2^-$ **<u>undergoes hydrolysis to produce OH^- ions</u>.**

$CH_3CH(OH)CO_2^-$(aq) + H_2O(l) ⇌ $CH_3CH(OH)CO_2H$(aq) + OH^-(aq)

The OH^- ions generated by the hydrolysis reaction causes the solution to be basic.

(f) See the graph in **(a)**.
[**Note:** Effective buffer range is p$K_a \pm 1$ i.e. from pH 2.9 to 4.9.]

7 **(a)** $C_6H_5CO_2H(aq) \rightleftharpoons C_6H_5CO_2^-(aq) + H^+(aq)$

(b) Let HA denote $C_6H_5CO_2H$.

$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$

At equilibrium, $[H^+] = [A^-]$
$[HA]_{eqm} \approx [HA]_{initial} = 0.010$ mol dm^{-3},
since benzoic acid is a weak acid with a small K_a

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[H^+]^2}{[HA]}$$

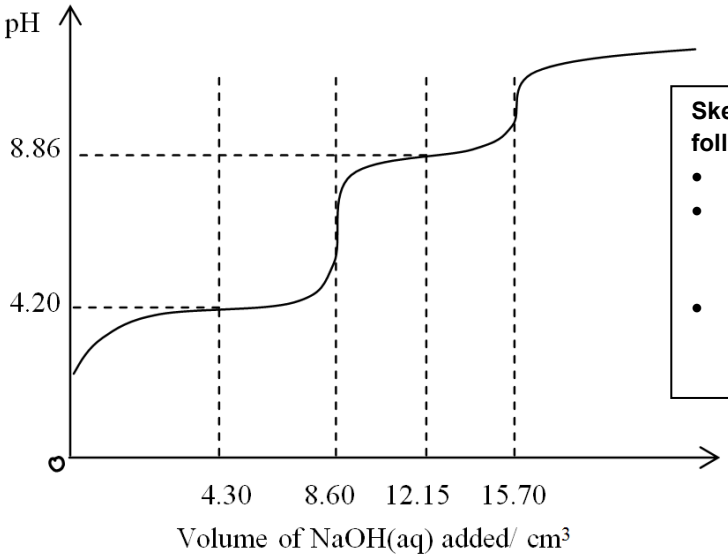
$$[H^+] = \sqrt{K_a \times [HA]} = \sqrt{10^{-4.20} \times 0.010} = 7.943 \times 10^{-4}\ \text{mol dm}^{-3}$$

$$\text{pH} = -\lg\left(7.943 \times 10^{-4}\right) = \underline{\mathbf{3.10}}$$

(c) Benzoic acid is a **stronger acid** than phenylboronic acid.

This is because benzoic acid has a **smaller pK_a** (or a larger K_a) than phenylboronic acid, which indicates that benzoic acid **dissociates to a greater extent** in aqueous solution.

(d) (i) A sketch of the titration curve.



Sketch should include the following:

- labelled axes
- the 2 pK_a values with the 2 corresponding volumes of NaOH
- volumes of NaOH required to reach the 2 equivalence points

(ii) When the first indicator changes colour, benzoic acid (the stronger acid) is completely neutralised.

$n_{\text{benzoic acid}}$ present $= n_{NaOH}$ used for neutralisation $= \dfrac{8.6}{1000} \times 0.050 = 4.30 \times 10^{-4}$ mol

$[\text{benzoic acid}]$ in **X** $= \dfrac{4.30 \times 10^{-4}}{10.0} \times 1000 =$ **$\underline{\text{0.0430 mol dm}^{-3}}$**

$n_{\text{phenylboronic acid}}$ present $= n_{NaOH}$ used for neutralisation $= \dfrac{7.1}{1000} \times 0.050 = 3.55 \times 10^{-4}$ mol

$[\text{phenylboronic acid}]$ in **X** $= \dfrac{3.55 \times 10^{-4}}{10.0} \times 1000 =$ **$\underline{\text{0.0355 mol dm}^{-3}}$**