MEGA LECTURE · A LEVEL CHEMISTRY

Chapter 22: Buffers and Titration Curves

OCR A Level Chemistry A (H432) — Structured Practice Worksheet

SPECIFICATION	TOTAL MARKS	SUGGESTED TIME
3.3.3 Acids, bases and buffers: buffer solutions and pH titration curves	84	1 hour 41 minutes

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Instructions to candidates

- Answer **all** questions in the spaces provided. Show all working in calculations.
- Marks are shown in brackets [ ] at the end of each part.

Question 1 — Buffer solutions: definition and composition 3.3.3

a. State what is meant by a *buffer solution*. **[2]**

b. State the two components required to make an acidic buffer solution, and give one example of a suitable combination. **[3]**

c. Explain why a mixture of a strong acid and its salt (e.g. HCl and NaCl) cannot act as a buffer solution. **[3]**

Question 1 total: 8 marks

Question 2 — How an acidic buffer works 3.3.3

A buffer solution contains ethanoic acid and sodium ethanoate: $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

a. Explain, using the equilibrium above, why this buffer solution contains a large "reservoir" of both CH_3COOH and CH_3COO^-. **[3]**

b. Explain what happens, in terms of the equilibrium, when a small amount of $H^+(aq)$ is added to this buffer. **[4]**

c. Explain what happens, in terms of the equilibrium, when a small amount of $OH^-(aq)$ is added to this buffer. **[3]**

Question 2 total: 10 marks

Question 3 — Calculating the pH of a buffer solution (I) 3.3.3

A buffer solution is prepared containing 0.400 mol dm^{-3} CH_3COOH and 0.250 mol dm^{-3} CH_3COONa. $K_a(CH_3COOH) = 1.74 \times 10^{-5}$ mol dm^{-3}.

a. Write the expression for K_a for CH_3COOH. **[2]**

b. State the two assumptions made when using this expression to calculate the pH of a buffer solution. **[2]**

c. Calculate the pH of this buffer solution. Show your working. **[6]**

Question 3 total: 10 marks

Question 4 — Preparing a buffer of a given pH

3.3.3

A student wants to prepare 1.00 dm^3 of a buffer solution with pH = 5.00, using 1.00 dm^3 of 0.500 mol dm^{-3} CH_3COOH solution and solid sodium ethanoate, CH_3COONa (M_r = 82.0). $K_a(CH_3COOH) = 1.74 \times 10^{-5}$ mol dm^{-3}.

a. Calculate the ratio $[CH_3COO^-]$ / $[CH_3COOH]$ required to give pH = 5.00. **[4]**

b. Given that the 1.00 dm^3 of CH_3COOH solution contains 0.500 mol CH_3COOH, calculate the number of moles of CH_3COONa required. **[3]**

c. Calculate the mass of solid sodium ethanoate that must be dissolved to prepare this buffer. **[3]**

Question 4 total: 10 marks

Question 5 — Calculating the pH of a buffer solution (II)

3.3.3

A buffer solution contains 0.300 mol CH_3COOH and 0.300 mol CH_3COONa dissolved in 1.00 dm^3 of solution. $K_a(CH_3COOH) = 1.74 \times 10^{-5}$ mol dm^{-3}.

a. Calculate the pH of this initial buffer solution. **[3]**

b. 0.0500 mol of HCl gas is bubbled into this buffer, and the volume remains 1.00 dm^3. Assuming the HCl reacts completely with CH_3COO^-, calculate the new pH of the buffer. **[6]**

c. State and explain how the pH change calculated in (b) compares with the pH change that would occur if the same amount of HCl were added to 1.00 dm^3 of pure water instead. **[3]**

Question 5 total: 12 marks

Question 6 — Biological importance of buffers 3.3.3

The carbonic acid/hydrogencarbonate buffer system helps maintain the pH of human blood: $H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$

a. State the approximate pH range within which human blood is normally maintained. **[1]**

b. Explain, using the equilibrium above, how this buffer system resists a fall in blood pH (e.g. from lactic acid produced during vigorous exercise). **[3]**

c. Explain why it is important that blood pH is maintained within a narrow range, and outline how the removal of CO_2 by the lungs helps regulate this buffer system. **[4]**

Question 6 total: 8 marks

Question 7 — pH titration curves 3.3.3

a. Sketch a graph of pH (y-axis) against volume of NaOH added (x-axis) for the titration of 25.0 cm^3 of 0.100 mol dm^{-3} HCl with 0.100 mol dm^{-3} NaOH. Label the initial pH, the equivalence point volume, and the equivalence point pH. **[3]**

Sketch: pH (y-axis) vs volume of NaOH added / cm³ (x-axis)

b. State the approximate pH at the equivalence point for: (i) a strong acid–strong base titration, (ii) a weak acid–strong base titration, (iii) a strong acid–weak base titration. **[3]**

c. Compare the size of the vertical (steep) region of the pH curve for a weak acid–strong base titration with that of a strong acid–strong base titration. **[3]**

d. Explain why a weak acid–weak base titration curve does not show a sharp vertical region (obvious point of inflection). **[3]**

Question 7 total: 12 marks

Question 8 — Choosing a suitable indicator

3.3.3

Indicator	pH range	Colour change (acid → base)
Methyl orange	3.1 – 4.4	red → yellow
Phenolphthalein	8.3 – 10.0	colourless → pink

a. Explain why phenolphthalein is suitable for both strong acid–strong base and weak acid–strong base titrations, but methyl orange is not suitable for the weak acid–strong base titration. **[3]**

b. Suggest which of these two indicators should be used for a strong acid–weak base titration, and explain your choice. **[2]**

c. Explain why no indicator is suitable for a weak acid–weak base titration. **[3]**

Question 8 total: 8 marks

Question 9 — Extended response

Synoptic

Explain, with reference to the expression $[H^+] = K_a \times [HA]/[A^-]$ and to the shape of a weak acid–strong base pH titration curve, why a buffer solution has its greatest capacity to resist pH change when it is prepared close to the pK_a of the weak acid used (i.e. when $[HA] \approx [A^-]$), and why buffer capacity decreases as either component becomes depleted, such as near the start or near the equivalence point of a titration curve. **[6]**

This question is marked using a level of response.

Question 9 total: 6 marks

PAPER TOTAL: 84 MARKS

MARK SCHEME — Chapter 22: Buffers and Titration Curves

Award marks independently unless stated. ECF applies to calculations.

Q1 (a) [2]

- A solution that resists (minimises/opposes) changes in pH (1) when small amounts of acid or base are added, or on dilution (1).

Q1 (b) [3]

- A weak acid (1) and a salt of its conjugate base / the salt of that weak acid with a strong base (1).
- Example: CH_3COOH and CH_3COONa (sodium ethanoate) — or any valid weak acid/salt pair. (1)

Q1 (c) [3]

- HCl is a strong acid and is fully/completely dissociated (1).
- So there is no reservoir of undissociated HA to release H^+ when OH^- is added / no equilibrium to shift to replace removed H^+ (1).
- Cl^- is not a base (conjugate base of a strong acid) and cannot react with/remove added H^+ (1).

Q2 (a) [3]

- The buffer contains a high concentration of undissociated CH_3COOH, from the weak acid (1).
- The buffer contains a high concentration of CH_3COO^-, mainly from the fully dissociated salt CH_3COONa (1).
- Both species are present in large, roughly comparable amounts compared with $[H^+]$ (1).

Q2 (b) [4]

- Added H^+ reacts with/combines with CH_3COO^- (1).
- The position of equilibrium shifts left, in the direction that removes the added H^+ (1).
- Forming more (undissociated) CH_3COOH (1).
- Since CH_3COO^- is present in large excess (reservoir), $[H^+]$ rises only very slightly, so pH is roughly maintained (1).

Q2 (c) [3]

- Added OH^- reacts with/removes H^+ (from CH_3COOH dissociation) (1).
- The equilibrium shifts right, and more CH_3COOH dissociates to replace the H^+ removed (1).
- Since CH_3COOH is present in large excess (reservoir), $[H^+]$ falls only slightly, so pH is roughly maintained (1).

Q3 (a) [2]

- $K_a = [H^+][CH_3COO^-] / [CH_3COOH]$ (2)

Q3 (b) [2]

- $[CH_3COO^-] \approx$ concentration of salt added (dissociation of the acid is suppressed/negligible due to the common ion effect) (1).
- $[CH_3COOH] \approx$ initial (stated) concentration of the acid, since very little dissociates (1).

Q3 (c) [6]

- $[H^+] = K_a \times [HA]/[A^-] = 1.74 \times 10^{-5} \times (0.400/0.250)$ (1)
- $= 1.74 \times 10^{-5} \times 1.60$ (1)
- $= 2.784 \times 10^{-5}$ mol dm^{-3} (1)
- $pH = -\log_{10}(2.784 \times 10^{-5})$ (1)
- $= 4.555$ (1)
- pH = 4.56 (2 d.p.) (1)

Q4 (a) [4]

- $[H^+] = 10^{-5.00} = 1.00 \times 10^{-5}$ mol dm^{-3} (1)
- $K_a = [H^+][A^-]/[HA]$, rearranging: $[A^-]/[HA] = K_a/[H^+]$ (1)
- $= 1.74 \times 10^{-5} \div 1.00 \times 10^{-5}$ (1)
- $= 1.74$ (1)

Q4 (b) [3]

- Moles CH_3COO^- required = ratio × moles CH_3COOH (1)
- = 1.74 × 0.500 (1)
- = 0.870 mol (1)

Q4 (c) [3]

- Mass = moles × M_r (1)
- = 0.870 × 82.0 (1)
- = 71.3 g (1)

Q5 (a) [3]

- Since moles of CH_3COOH and CH_3COONa are equal, [HA] = [A^-], so [H^+] = K_a = 1.74×10^{-5} mol dm^{-3} (1)
- pH = $-\log_{10}(1.74 \times 10^{-5})$ (1)
- pH = 4.76 (1)

Q5 (b) [6]

- New moles CH_3COO^- = 0.300 − 0.0500 = 0.250 mol (1)
- New moles CH_3COOH = 0.300 + 0.0500 = 0.350 mol (1)
- [H^+] = K_a × [HA]/[A^-] = 1.74×10^{-5} × (0.350/0.250) (1)
- = $1.74 \times 10^{-5} \times 1.40 = 2.436 \times 10^{-5}$ mol dm^{-3} (1)
- pH = $-\log_{10}(2.436 \times 10^{-5})$ (1)
- pH = 4.61 (1)

ECF: full marks for a correctly-calculated pH from an incorrect ratio, provided the method is correct.

Q5 (c) [3]

- In pure water, adding 0.0500 mol HCl to 1.00 dm^3 gives [H^+] ≈ 0.0500 mol dm^{-3}, so pH would fall drastically, from 7 to about 1.3 (1).
- In the buffer, the pH only falls slightly, from 4.76 to 4.61 (1).
- Because the buffer's CH_3COO^- reservoir reacts with/mops up the added H^+, resisting the pH change, whereas pure water has no such reservoir (1).

Q6 (a) [1]

- pH 7.35 – 7.45 (accept "approximately 7.4") (1)

Q6 (b) [3]

- Extra H^+ (from lactic acid) reacts with HCO_3^- (1).
- The equilibrium shifts left, forming more H_2CO_3 (1).
- This removes the added H^+ from solution, so pH does not fall significantly (1).

Q6 (c) [4]

- Enzymes/proteins in the body are sensitive to pH and can denature/lose function outside a narrow pH range (1).
- If blood pH falls too far (acidosis) or rises too far (alkalosis), this can be life-threatening (1).
- H_2CO_3 is in equilibrium with dissolved CO_2 ($CO_2 + H_2O \rightleftharpoons H_2CO_3$), and CO_2 is removed from the body by exhalation via the lungs (1).
- Increasing the breathing rate removes CO_2 faster, shifting the equilibrium to remove H_2CO_3/H^+, helping to raise blood pH back towards normal (1).

Q7 (a) [3]

- Curve starts at a low pH (approximately pH 1) (1).
- Large/steep vertical region around 25.0 cm^3 (the equivalence point), climbing sharply from about pH 3 to pH 11 (1).
- Curve levels off at a high pH (approximately pH 13) once excess NaOH has been added (1).

Q7 (b) [3]

- (i) Strong acid–strong base: pH = 7 (1).
- (ii) Weak acid–strong base: pH > 7 (approximately 8–9) (1).
- (iii) Strong acid–weak base: pH < 7 (approximately 5–6) (1).

Q7 (c) [3]

- The vertical region for a weak acid–strong base titration is much smaller/shorter (1).
- E.g. approximately pH 7–10, rather than approximately pH 3–11 for strong acid–strong base (1).
- Because the weak acid and its conjugate base act as a buffer, resisting pH change before the equivalence point is reached (1).

Q7 (d) [3]

- Neither reactant nor product is fully dissociated — both the weak acid/its conjugate base and the weak base/its conjugate acid are present in solution, forming buffer regions (1).
- The pH changes gradually/rises steadily throughout the titration rather than jumping sharply near the equivalence point (1).
- The equivalence point is therefore difficult to identify accurately from the curve, or using an indicator (1).

Q8 (a) [3]

- The vertical region of a strong acid–strong base curve spans roughly pH 3–11, and of a weak acid–strong base curve spans roughly pH 7–10 — both include phenolphthalein's range (8.3–10.0) (1).
- Methyl orange's range (3.1–4.4) falls within the strong acid–strong base vertical region, but is below/outside the weak acid–strong base vertical region (which starts around pH 7) (1).
- So methyl orange would change colour gradually, well before the equivalence point, in the weak acid–strong base titration, rather than sharply at it (1).

Q8 (b) [2]

- Methyl orange (1).
- Because the vertical region for a strong acid–weak base titration spans roughly pH 3–7, which includes methyl orange's range but not phenolphthalein's (1).

Q8 (c) [3]

- There is no sharp vertical region/no sudden change in pH at the equivalence point (1).
- The pH changes gradually over a wide volume range (1).
- So no single indicator changes colour sharply enough, or close enough to the equivalence point volume, to give a distinct, reliable end point (1).

Q9 [6] — Level of response

- **Level 3 (5–6):** Explains fully and correctly both why buffer capacity is greatest when [HA] ≈ [A^-] (pH = pK_a) and why capacity falls away from this point, linking clearly to the $[H^+] = K_a[HA]/[A^-]$ expression and to the titration curve.
- **Level 2 (3–4):** Explains one of the two ideas (best buffering at pH = pK_a, or reduced capacity when a component is depleted) well, with a partial or less complete explanation of the other.
- **Level 1 (1–2):** States that buffering capacity varies, with limited or generic reasoning; little reference to the K_a expression or the titration curve.

Indicative content: *When [HA] = [A^-], the ratio [HA]/[A^-] = 1, so [H^+] = K_a and pH = pK_a; at this point both the "reservoir" of HA (to mop up added OH^-) and the reservoir of A^- (to mop up added H^+) are present in their largest, most balanced amounts, so small additions of acid or base change the ratio [HA]/[A^-] only slightly and pH is well resisted. As the buffer composition moves away from this point (e.g. one component becomes small relative to the other), a small addition of acid or base causes a proportionally much larger change in the ratio [HA]/[A^-], so pH changes more. On a weak acid–strong base titration curve, the half-equivalence point (where exactly half the acid has been converted to its conjugate base) corresponds to [HA] = [A^-] and pH = pK_a; the curve is flattest (most resistant to pH change, i.e. best buffering) around this point, and becomes steeper towards the start (little A^- present) and especially near the equivalence point (almost no HA remaining), where buffering capacity is essentially lost.*

END OF MARK SCHEME · 84 marks