AQA A2 Chemistry Answer Sheet – Acids and Bases - pH and Buffer Solutions

Model Answers and Mark Schemes | Total Marks: 45

Question 1 (1 mark)

> *Define the term pH.*

MODEL ANSWER

pH is the negative logarithm to the base 10 of the hydrogen ion concentration.

MARK SCHEME

- ✓ Negative logarithm to base 10 of hydrogen ion concentration / $-\log_{10}[H^+]$ [1 mark]

Question 2 (2 marks)

> *Explain why the pH of a strong acid solution can be calculated directly from its concentration, whereas for a weak acid, an equilibrium constant is required.*

MODEL ANSWER

Strong acids fully dissociate in solution, so the concentration of H^+ ions is equal to the initial concentration of the acid. Weak acids only partially dissociate, establishing an equilibrium, so an equilibrium constant (K_a) is needed to determine the $[H^+]$ at equilibrium.

MARK SCHEME

- ✓ Strong acids fully dissociate / ionise [1 mark]
- ✓ Weak acids partially dissociate / ionise / establish equilibrium, so K_a needed to find $[H^+]$ [1 mark]

Question 3 (4 marks)

Calculate the pH of a 0.025 mol dm^{-3} solution of nitric acid (HNO_3).

M O D E L   A N S W E R

HNO_3 is a strong acid, so it fully dissociates:

$HNO_3 \rightarrow H^+ + NO_3^-$

$[H^+] = [HNO_3] = 0.025$ mol dm^{-3}

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

$pH = -\log_{10}(0.025)$

$pH = 1.60$

S T E P - B Y - S T E P   W O R K I N G

Step 1: Recognise HNO_3 is a strong acid and fully dissociates, so $[H^+] = [HNO_3]$.

Step 2: State the formula for pH.

Step 3: Substitute the $[H^+]$ value into the pH formula.

Step 4: Calculate the final pH, rounding to an appropriate number of decimal places (usually 2).

M A R K   S C H E M E

- ✓ $[H^+] = 0.025$ mol dm^{-3} (due to full dissociation) [1 mark]
- ✓ $pH = -\log_{10}[H^+]$ [1 mark]
- ✓ $pH = -\log_{10}(0.025)$ [1 mark]
- ✓ pH = 1.60 (2 dp) [1 mark]

Question 4 (5 marks)

A student prepares a 0.100 mol dm^{-3} solution of ethanoic acid (CH_3COOH). The K_a for ethanoic acid is 1.75 × 10^{-5} mol dm^{-3} at 298 K.

Calculate the pH of this ethanoic acid solution.

MODEL ANSWER

$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$

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

Assume $[H^+] = [CH_3COO^-]$ and $[CH_3COOH]_{initial} \approx [CH_3COOH]_{equilibrium}$

$K_a = [H^+]^2 / [CH_3COOH]$

$[H^+]^2 = K_a \times [CH_3COOH]$

$[H^+]^2 = (1.75 \times 10^{-5}) \times 0.100$

$[H^+]^2 = 1.75 \times 10^{-6}$

$[H^+] = \sqrt{(1.75 \times 10^{-6})} = 1.3228 \times 10^{-3}$ mol dm^{-3}

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

$pH = -\log_{10}(1.3228 \times 10^{-3})$

pH = 2.88

STEP-BY-STEP WORKING

Step 1: Write the K_a expression for ethanoic acid.

Step 2: Make the standard assumptions for a weak acid calculation: $[H^+] = [CH_3COO^-]$ and the equilibrium concentration of the acid is approximately equal to its initial concentration.

Step 3: Rearrange the K_a expression to solve for $[H^+]^2$ or $[H^+]$.

Step 4: Substitute the given values for K_a and $[CH_3COOH]$ and calculate $[H^+]$.

Step 5: Use $pH = -\log_{10}[H^+]$ to find the pH, rounding to 2 decimal places.

MARK SCHEME

- ✓ K_a expression (or implied by working) [1 mark]
- ✓ Assumption $[H^+] = [CH_3COO^-]$ and $[CH_3COOH]_{initial} \approx [CH_3COOH]_{equilibrium}$ [1 mark]
- ✓ $[H^+]^2 = K_a \times [CH_3COOH]$ OR $[H^+] = \sqrt{(K_a \times [CH_3COOH])}$ [1 mark]
- ✓ $[H^+] = 1.32 \times 10^{-3}$ mol dm^{-3} [1 mark]
- ✓ pH = 2.88 (2 dp) [1 mark]

Question 5 (1 mark)

> *State the meaning of the term buffer solution.*

MODEL ANSWER

A buffer solution is a solution that resists changes in pH when small amounts of acid or alkali are added.

MARK SCHEME

✓ Resists changes in pH when small amounts of acid/alkali are added [1 mark]

Question 6 (3 marks)

> *Explain how a buffer solution made from a weak acid and its conjugate base resists a decrease in pH when a small amount of strong acid is added.*

MODEL ANSWER

When a small amount of strong acid (H^+) is added, the H^+ ions react with the conjugate base (A^-) present in the buffer solution.

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

This removes most of the added H^+ ions, shifting the equilibrium to the right and preventing a significant increase in $[H^+]$, thus resisting a decrease in pH.

MARK SCHEME

✓ Added H^+ reacts with the conjugate base (A^-) [1 mark]

✓ Equation: $A^- + H^+ \rightarrow HA$ (or specific example) [1 mark]

✓ Removes added H^+ / shifts equilibrium to the right, preventing significant change in pH [1 mark]

Question 7 (4 marks)

A buffer solution is prepared by mixing 25.0 cm^3 of 0.200 mol dm^{-3} propanoic acid (CH_3CH_2COOH) with 25.0 cm^3 of 0.100 mol dm^{-3} sodium propanoate (CH_3CH_2COONa). The K_a for propanoic acid is 1.30 × 10^{-5} mol dm^{-3}.

Calculate the pH of this buffer solution.

M O D E L   A N S W E R

Total volume = 25.0 + 25.0 = 50.0 cm^3 = 0.050 dm^3
Moles of propanoic acid = 0.200 mol dm^{-3} × 0.025 dm^3 = 0.00500 mol
Moles of propanoate = 0.100 mol dm^{-3} × 0.025 dm^3 = 0.00250 mol

Concentration of propanoic acid = 0.00500 mol / 0.050 dm^3 = 0.100 mol dm^{-3}
Concentration of propanoate = 0.00250 mol / 0.050 dm^3 = 0.0500 mol dm^{-3}

Using the Henderson-Hasselbalch equation: $pH = pK_a + \log_{10}([A^-]/[HA])$
$pK_a = -\log_{10}(1.30 \times 10^{-5}) = 4.886$
$pH = 4.886 + \log_{10}(0.0500 / 0.100)$
$pH = 4.886 + \log_{10}(0.5)$
$pH = 4.886 - 0.301$
$pH = 4.585 \approx 4.59$

S T E P - B Y - S T E P   W O R K I N G

Step 1: Calculate the initial moles of the weak acid and its conjugate base.
Step 2: Calculate the new concentrations of the weak acid and conjugate base in the total volume of the buffer solution. Alternatively, use moles directly in the Henderson-Hasselbalch equation if the volume cancels out.
Step 3: Calculate pK_a from the given K_a value.
Step 4: Substitute the concentrations (or moles) and pK_a into the Henderson-Hasselbalch equation ($pH = pK_a + \log_{10}([A^-]/[HA])$) or use the K_a expression to find $[H^+]$ and then pH.
Step 5: Calculate the final pH, rounding to 2 decimal places.

M A R K   S C H E M E

- ✓ Calculates moles of acid and conjugate base (or concentrations in mixed volume) [1 mark]
- ✓ Correctly calculates pK_a [1 mark]
- ✓ Correctly substitutes values into Henderson-Hasselbalch equation or K_a expression [1 mark]
- ✓ pH = 4.59 (2 dp) [1 mark]

Question 8 (2 marks)

> *Explain why a buffer solution has a limited capacity to resist changes in pH.*

MODEL ANSWER

A buffer solution has a limited capacity because it contains a finite amount of the weak acid and its conjugate base. Once these components are used up by the addition of significant amounts of strong acid or alkali, the buffer can no longer resist changes in pH.

MARK SCHEME

- ✓ Finite amount of weak acid and conjugate base [1 mark]
- ✓ Once these are used up, the buffer can no longer resist pH change [1 mark]

Question 9 (1 mark)

> *Write the expression for the ionic product of water, K_w.*

MODEL ANSWER

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

MARK SCHEME

- ✓ $K_w = [H^+][OH^-]$ [1 mark]

Question 10 (3 marks)

At 373 K, $K_w = 5.47 \times 10^{-14}$ mol^2 dm^{-6}. Calculate the pH of pure water at this temperature.

M O D E L   A N S W E R

In pure water, $[H^+] = [OH^-]$

$K_w = [H^+]^2$

$[H^+] = \sqrt{K_w}$

$[H^+] = \sqrt{(5.47 \times 10^{-14})}$

$[H^+] = 2.338 \times 10^{-7}$ mol dm^{-3}

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

$pH = -\log_{10}(2.338 \times 10^{-7})$

pH = 6.63

S T E P - B Y - S T E P   W O R K I N G

Step 1: Recognise that in pure water, $[H^+] = [OH^-]$, so $K_w = [H^+]^2$.

Step 2: Calculate $[H^+]$ by taking the square root of K_w.

Step 3: Calculate pH using $pH = -\log_{10}[H^+]$, rounding to 2 decimal places.

M A R K   S C H E M E

✓ $[H^+] = \sqrt{K_w}$ (or implied) [1 mark]

✓ $[H^+] = 2.34 \times 10^{-7}$ mol dm^{-3} [1 mark]

✓ pH = 6.63 (2 dp) [1 mark]

Question 11 (5 marks)

A student titrates 25.0 cm^3 of 0.150 mol dm^{-3} ethanoic acid (CH_3COOH) with 0.100 mol dm^{-3} sodium hydroxide (NaOH). The K_a for ethanoic acid is 1.75×10^{-5} mol dm^{-3}.

Calculate the pH of the solution after 10.0 cm^3 of NaOH has been added.

MODEL ANSWER

Initial moles of CH_3COOH = 0.150 mol dm^{-3} × 0.025 dm^3 = 0.00375 mol
Moles of NaOH added = 0.100 mol dm^{-3} × 0.010 dm^3 = 0.00100 mol

Reaction: $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
Moles of CH_3COOH remaining = 0.00375 - 0.00100 = 0.00275 mol
Moles of CH_3COONa formed = 0.00100 mol

Total volume = 25.0 cm^3 + 10.0 cm^3 = 35.0 cm^3 = 0.035 dm^3
$[CH_3COOH]$ = 0.00275 mol / 0.035 dm^3 = 0.07857 mol dm^{-3}
$[CH_3COONa]$ = 0.00100 mol / 0.035 dm^3 = 0.02857 mol dm^{-3}

Using $K_a = [H^+][CH_3COO^-] / [CH_3COOH]$
$[H^+] = K_a \times [CH_3COOH] / [CH_3COO^-]$
$[H^+] = (1.75 \times 10^{-5}) \times (0.07857 / 0.02857)$
$[H^+] = (1.75 \times 10^{-5}) \times 2.749$
$[H^+] = 4.811 \times 10^{-5}$ mol dm^{-3}
$pH = -\log_{10}(4.811 \times 10^{-5})$
pH = 4.32

STEP-BY-STEP WORKING

Step 1: Calculate the initial moles of the weak acid.
Step 2: Calculate the moles of strong base added.
Step 3: Determine the moles of weak acid remaining and the moles of conjugate base (salt) formed after the reaction.
Step 4: Calculate the total volume of the solution.
Step 5: Calculate the concentrations of the remaining weak acid and the formed conjugate base in the total volume.
Step 6: Use the K_a expression ($K_a = [H^+][A^-]/[HA]$) to calculate $[H^+]$.
Step 7: Calculate the pH using $pH = -\log_{10}[H^+]$, rounding to 2 decimal places.

MARK SCHEME

- ✓ Calculates initial moles of acid and moles of NaOH added [1 mark]
- ✓ Calculates moles of remaining acid and moles of salt formed [1 mark]
- ✓ Calculates concentrations of remaining acid and salt in total volume [1 mark]
- ✓ Correctly uses K_a expression to find $[H^+]$ [1 mark]
- ✓ pH = 4.32 (2 dp) [1 mark]

Question 12 (6 marks)

Consider the titration of 20.0 cm^3 of 0.100 mol dm^{-3} ammonia (NH_3) with 0.100 mol dm^{-3} hydrochloric acid (HCl). The K_b for ammonia is 1.80×10^{-5} mol dm^{-3}.

(a) [3 marks]

Calculate the pH of the ammonia solution before any HCl is added.

MODEL ANSWER

See model answer above.

MARK SCHEME

- ✓ See marking points [3 marks]

(b) [3 marks]

Calculate the pH at the equivalence point of this titration.

MODEL ANSWER

See model answer above.

MARK SCHEME

- ✓ See marking points [3 marks]

Question 13 (8 marks)

A buffer solution is prepared by dissolving 12.0 g of sodium ethanoate (CH_3COONa, M_r = 82.0) in 250 cm^3 of 0.500 mol dm^{-3} ethanoic acid (CH_3COOH). The K_a for ethanoic acid is 1.75 × 10^{-5} mol dm^{-3}.

(a) [4 marks]

Calculate the pH of this buffer solution.

MODEL ANSWER

See model answer above.

MARK SCHEME

✓ See marking points [4 marks]

(b) [4 marks]

A small amount of 0.050 mol dm^{-3} HCl is added to 100 cm^3 of the buffer solution prepared in part (a). Calculate the new pH of the buffer solution after 5.0 cm^3 of HCl has been added. Assume the volume change is negligible for the buffer components.

MODEL ANSWER

See model answer above.

MARK SCHEME

✓ See marking points [4 marks]