MEGA LECTURE · A LEVEL CHEMISTRY

Chapter 21: Acids, Bases and pH

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

SPECIFICATION	TOTAL MARKS	SUGGESTED TIME
3.3.3 Acids, bases and pH: Bronsted–Lowry theory, K_a, pK_a, K_w, pH calculations	92	1 hour 50 minutes

www.MegaLecture.com · WhatsApp: +92 323 509 4443

MEGA LECTURE · A LEVEL CHEMISTRY

Chapter 21: Acids, Bases and pH

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

SPECIFICATION	TOTAL MARKS	SUGGESTED TIME
3.3.3 Acids, bases and pH: Bronsted–Lowry theory, K_a, pK_a, K_w, pH calculations	92	1 hour 50 minutes

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.
- Unless otherwise stated, assume a temperature of 298 K, where $K_w = 1.00 \times 10^{-14}$ mol^2 dm^{-6}.

Question 1 — Bronsted–Lowry acids and bases 3.3.3

a. State the Bronsted–Lowry definitions of an acid and of a base. **[2]**

b. State what is meant by a *conjugate acid–base pair*. **[2]**

c. For the reaction $HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$, identify the Bronsted–Lowry acid, the base, and one conjugate acid–base pair. **[3]**

d. Write an equation for $NH_3(aq)$ acting as a Bronsted–Lowry base with water, and identify the conjugate acid formed. **[2]**

Question 1 total: 9 marks

Question 2 — Strong and weak acids and bases 3.3.3

a. State what is meant by a *strong acid*, in terms of dissociation. **[2]**

b. State what is meant by a *weak acid*, in terms of dissociation. **[2]**

c. Explain, in terms of degree of dissociation, why a 0.10 mol dm^{-3} solution of a strong acid has a much lower pH than a 0.10 mol dm^{-3} solution of a weak acid. **[3]**

d. Classify each of the following as a strong or weak acid: HCl, CH_3COOH, HNO_3, HF. **[2]**

Acid	Strong or weak?
HCl	
CH_3COOH	
HNO_3	
HF	

Question 2 total: 9 marks

Question 3 — Ka and pKa

3.3.3

a. Write the expression for the acid dissociation constant, K_a, for a weak acid HA, and state what it measures. **[2]**

b. State the equation linking pK_a and K_a. **[2]**

c. Ethanoic acid has $K_a = 1.74 \times 10^{-5}$ mol dm^{-3} at 298 K. Calculate its pK_a. **[2]**

d. A weak acid has $pK_a = 3.14$ at 298 K. Calculate its K_a, including units. **[2]**

e. Explain why a smaller value of pK_a corresponds to a stronger (weak) acid. **[2]**

Question 3 total: 10 marks

Question 4 — Kw and pH

3.3.3

a. Write the expression for K_w, and state its value at 298 K, including units. **[2]**

b. Write an equation for the equilibrium in water that gives rise to K_w. **[2]**

c. Define pH. **[2]**

d. Calculate $[H^+]$ for a solution with pH = 2.30. **[2]**

e. Calculate the pH of a solution with $[H^+] = 3.16 \times 10^{-5}$ mol dm^{-3}. **[2]**

Question 4 total: 10 marks

Question 5 — pH of strong acids

3.3.3

a. State the assumption made when calculating the pH of a strong, monoprotic acid from its concentration. **[2]**

b. Calculate the pH of 0.0500 mol dm^{-3} HCl(aq). **[3]**

c. Calculate the pH of 2.50×10^{-3} mol dm^{-3} HNO_3(aq). **[3]**

d. H_2SO_4 is a strong diprotic acid, fully dissociating to give two moles of H^+ per mole of acid. Calculate the concentration of H_2SO_4(aq) required to give a solution of pH = 1.00. **[3]**

Question 5 total: 11 marks

Question 6 — pH of strong bases 3.3.3

a. State the assumption made when calculating [OH^-] for a strong, monobasic base from its concentration. **[2]**

b. Calculate the pH of 0.0200 mol dm^{-3} NaOH(aq). **[4]**

c. $Ca(OH)_2$ fully dissociates, giving two moles of OH^- per mole of $Ca(OH)_2$. Calculate the pH of 5.00×10^{-2} mol dm^{-3} $Ca(OH)_2$(aq). **[4]**

Question 6 total: 10 marks

Question 7 — pH of weak acids

3.3.3

a. The pH of a weak acid may be estimated using $[H^+] = \sqrt{(K_a[HA])}$. State the two assumptions made in using this approximation. **[3]**

b. Calculate the pH of 0.200 mol dm^{-3} $CH_3COOH(aq)$, given $K_a = 1.74 \times 10^{-5}$ mol dm^{-3}. **[5]**

c. Calculate the pH of 0.0500 mol dm^{-3} $HF(aq)$, given $K_a = 5.6 \times 10^{-4}$ mol dm^{-3}. **[5]**

Question 7 total: 13 marks

Question 8 — Calculating Ka from pH

3.3.3

a. Rearrange $[H^+] = \sqrt{(K_a[HA])}$ to make K_a the subject. **[2]**

b. A 0.150 mol dm^{-3} solution of a weak acid HX has a pH of 2.95 at 298 K. Calculate the value of K_a for HX, including units, and hence its pK_a. **[5]**

c. Using your answer to (b), calculate the percentage of HX that has dissociated, and comment on whether the assumption that dissociation is small was valid for this acid. **[3]**

Question 8 total: 10 marks

Question 9 — Limitations of the weak acid approximation 3.3.3

a. State the two assumptions made in the approximation $[H^+] = \sqrt{(K_a[HA])}$, and identify a situation in which this approximation is likely to be inaccurate. **[2]**

b. Explain why, for a very dilute strong acid (e.g. 1×10^{-8} mol dm^{-3}), the assumption that $[H^+]$ from the dissociation of water is negligible breaks down, and state the effect this has on the actual pH compared with a naive calculation. **[2]**

Question 9 total: 4 marks

Question 10 — Extended response Synoptic

A student calculates the pH of a 0.0100 mol dm^{-3} solution of a weak acid, HA ($K_a = 6.3 \times 10^{-5}$ mol dm^{-3}), using the approximation $[H^+] = \sqrt{(K_a[HA])}$, and obtains a pH of 3.10. Discuss, with reference to the assumptions underlying this approximation, whether this calculated pH is likely to be a reliable estimate of the true pH of the solution. **[6]**

This question is marked using a level of response.

Question 10 total: 6 marks

PAPER TOTAL: 92 MARKS

MARK SCHEME — Chapter 21: Acids, Bases and pH

Award marks independently unless stated. ECF applies to calculations.

Q1 (a) [2]

- A Bronsted–Lowry acid is a proton (H^+) donor. (1)
- A Bronsted–Lowry base is a proton (H^+) acceptor. (1)

Q1 (b) [2]

- A conjugate acid–base pair consists of two species that differ (interconvert) by the transfer of a single proton (H^+). (1)
- The conjugate base is formed when the acid donates its proton; the conjugate acid is formed when the base accepts a proton. (1)

Q1 (c) [3]

- Acid: HCl (donates a proton to water). (1)
- Base: H_2O (accepts a proton from HCl). (1)
- Conjugate pair: HCl/Cl^-, or H_2O/H_3O^+ (either pair accepted). (1)

Q1 (d) [2]

- $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$. (1)
- Conjugate acid formed: NH_4^+. (1)

Q2 (a) [2]

- A strong acid fully (completely/100%) dissociates (ionises) in aqueous solution. (2)

Q2 (b) [2]

- A weak acid only partially (incompletely) dissociates in aqueous solution — an equilibrium exists between the undissociated acid and its ions. (2)

Q2 (c) [3]

- The strong acid fully dissociates, so $[H^+] \approx$ the full initial concentration of the acid, 0.10 mol dm^{-3}. (1)
- The weak acid only partially dissociates, so $[H^+]$ is much lower than 0.10 mol dm^{-3}. (1)
- Since $pH = -\log_{10}[H^+]$, the much higher $[H^+]$ of the strong acid gives it a much lower pH than the weak acid at the same concentration. (1)

Q2 (d) [2]

- HCl: strong; CH_3COOH: weak. (1)
- HNO_3: strong; HF: weak. (1)

Q3 (a) [2]

- $K_a = [H^+][A^-] / [HA]$. (1)
- K_a is a measure of the extent of dissociation (strength) of a weak acid — the larger the K_a, the stronger the (weak) acid. (1)

Q3 (b) [2]

- $pK_a = -\log_{10}K_a$ (2)

Q3 (c) [2]

- $pK_a = -\log_{10}(1.74 \times 10^{-5})$. (1)
- $pK_a = 4.76$. (1)

Q3 (d) [2]

- $K_a = 10^{-3.14}$. (1)
- $K_a = 7.24 \times 10^{-4}$ mol dm^{-3}. (1)

Q3 (e) [2]

- $pK_a = -\log_{10}K_a$, so a smaller pK_a corresponds to a larger K_a. (1)
- A larger K_a means the acid dissociates more (has a greater equilibrium $[H^+]$ for the same concentration), i.e. it is a stronger acid. (1)

Q4 (a) [2]

- $K_w = [H^+][OH^-]$. (1)
- $K_w = 1.00 \times 10^{-14}$ mol^2 dm^{-6} at 298 K. (1)

Q4 (b) [2]

- $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$ (or $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$). (2)

Q4 (c) [2]

- $pH = -\log_{10}[H^+]$, where $[H^+]$ is in mol dm^{-3}. (2)

Q4 (d) [2]

- $[H^+] = 10^{-2.30}$. (1)
- $[H^+] = 5.01 \times 10^{-3}$ mol dm^{-3}. (1)

Q4 (e) [2]

- $pH = -\log_{10}(3.16 \times 10^{-5})$. (1)
- pH = 4.50. (1)

Q5 (a) [2]

- A strong monoprotic acid is assumed to fully (100%) dissociate. (1)
- So $[H^+]$ at equilibrium is taken to equal the initial concentration of the acid. (1)

Q5 (b) [3]

- $[H^+] = 0.0500$ mol dm^{-3} (full dissociation). (1)
- $pH = -\log_{10}(0.0500)$. (1)
- pH = 1.30. (1)

Q5 (c) [3]

- $[H^+] = 2.50 \times 10^{-3}$ mol dm^{-3}. (1)
- $pH = -\log_{10}(2.50 \times 10^{-3})$. (1)
- pH = 2.60. (1)

Q5 (d) [3]

- $[H^+] = 10^{-1.00} = 0.100$ mol dm^{-3}. (1)
- H_2SO_4 gives 2 mol H^+ per mole of acid, so $[H_2SO_4] = [H^+] \div 2$. (1)
- $[H_2SO_4] = 0.100 \div 2 = 0.0500$ mol dm^{-3}. (1)

Q6 (a) [2]

- A strong, monobasic base is assumed to fully dissociate. (1)
- So $[OH^-]$ at equilibrium is taken to equal the initial concentration of the base. (1)

Q6 (b) [4]

- $[OH^-] = 0.0200$ mol dm^{-3} (full dissociation). (1)
- $[H^+] = K_w \div [OH^-] = (1.00 \times 10^{-14}) \div 0.0200$. (1)
- $[H^+] = 5.00 \times 10^{-13}$ mol dm^{-3}. (1)
- $pH = -\log_{10}(5.00 \times 10^{-13}) = 12.30$. (1)

Q6 (c) [4]

- $[OH^-] = 2 \times 5.00 \times 10^{-2} = 0.100$ mol dm^{-3}. (1)
- $[H^+] = K_w \div [OH^-] = (1.00 \times 10^{-14}) \div 0.100$. (1)
- $[H^+] = 1.00 \times 10^{-13}$ mol dm^{-3}. (1)
- $pH = -\log_{10}(1.00 \times 10^{-13}) = 13.00$. (1)

Q7 (a) [3]

- The extent of dissociation of the weak acid is assumed to be small, so [HA] at equilibrium ≈ [HA] initial (undissociated concentration is not significantly reduced). (1)
- $[H^+]$ from the dissociation of water is assumed to be negligible compared with $[H^+]$ produced by the acid. (1)
- These assumptions allow $[H^+] = [A^-]$ and $[HA]_{eqm} \approx [HA]_{initial}$ to be used in the K_a expression. (1)

Q7 (b) [5]

- $[H^+] = \sqrt{(K_a[HA])} = \sqrt{(1.74 \times 10^{-5} \times 0.200)}$. (1)
- $= \sqrt{(3.48 \times 10^{-6})}$. (1)
- $[H^+] = 1.87 \times 10^{-3}$ mol dm^{-3}. (1)
- $pH = -\log_{10}(1.87 \times 10^{-3})$. (1)
- pH = 2.73. (1)

Q7 (c) [5]

- $[H^+] = \sqrt{(K_a[HA])} = \sqrt{(5.6 \times 10^{-4} \times 0.0500)}$. (1)
- $= \sqrt{(2.8 \times 10^{-5})}$. (1)
- $[H^+] = 5.29 \times 10^{-3}$ mol dm^{-3}. (1)
- $pH = -\log_{10}(5.29 \times 10^{-3})$. (1)
- pH = 2.28. (1)

Q8 (a) [2]

- $[H^+]^2 = K_a[HA]$. (1)
- $K_a = [H^+]^2 / [HA]$. (1)

Q8 (b) [5]

- $[H^+] = 10^{-2.95} = 1.12 \times 10^{-3}$ mol dm^{-3}. (1)
- $K_a = [H^+]^2 / [HA] = (1.12 \times 10^{-3})^2 \div 0.150$. (1)
- $K_a = (1.26 \times 10^{-6}) \div 0.150 = 8.39 \times 10^{-6}$ mol dm^{-3}. (1)
- $pK_a = -\log_{10}(8.39 \times 10^{-6})$. (1)
- $pK_a = 5.08$. (1)

Q8 (c) [3]

- % dissociated = $([H^+] \div [HA]_{initial}) \times 100 = (1.12 \times 10^{-3} \div 0.150) \times 100$. (1)
- = 0.75%. (1)
- Since this is much less than 5%, the assumption that dissociation is small (negligible) compared with the initial concentration is valid for this acid. (1)

Q9 (a) [2]

- Assumptions: dissociation of the weak acid is small (so $[HA]_{eqm} \approx [HA]_{initial}$), and $[H^+]$ from water is negligible compared with $[H^+]$ from the acid. (1)
- The approximation is likely to be inaccurate for very dilute weak acid solutions, or weak acids with a relatively large K_a (where dissociation is not small). (1)

Q9 (b) [2]

- At very low acid concentration, the $[H^+]$ produced by the acid itself becomes comparable to (similar in size to) the $[H^+]$ produced by the autoionisation of water, so it is no longer negligible. (1)
- The actual pH is higher (less acidic, closer to pH 7) than predicted by the simple calculation, which would otherwise (incorrectly) predict the pH to keep falling indefinitely as concentration decreases. (1)

Q10 [6] — Level of response

- **Level 3 (5–6):** Discusses both assumptions clearly, correctly calculates or estimates the percentage dissociation (≈ 7.9%) and uses it to make a well-reasoned judgement about the reliability of the approximation, with a balanced, logical conclusion.
- **Level 2 (3–4):** Discusses one assumption in good detail (e.g. dissociation being small), with limited or no reference to the other assumption or to a supporting calculation.
- **Level 1 (1–2):** States that the approximation may not be reliable with limited or generic reasoning; little reference to the specific assumptions or to any supporting calculation.

Indicative content: *The approximation $[H^+] = \sqrt{(K_a[HA])}$ relies on two assumptions: that the degree of dissociation of the acid is small compared with its initial concentration (so $[HA]_{eqm} \approx [HA]_{initial}$), and that $[H^+]$ from the autoionisation of water is negligible compared with $[H^+]$ from the acid. For this acid, $[H^+] = \sqrt{(6.3 \times 10^{-5} \times 0.0100)} = 7.94 \times 10^{-4}$ mol dm^{-3}, giving the stated pH of 3.10. The percentage dissociation is $(7.94 \times 10^{-4} \div 0.0100) \times 100 \approx 7.9\%$, which is above the commonly used 5% threshold for the approximation to be considered reliable. This suggests the assumption that dissociation is small is only borderline valid here, so the calculated pH is a reasonable estimate but may carry a small but non-negligible error; a more rigorous (quadratic) treatment of the equilibrium would give a slightly more accurate value. The assumption*

regarding water's contribution to $[H^+]$ remains valid here, since $[H^+]$ from the acid ($\approx 10^{-4}$ mol dm^{-3}) is far greater than that from water alone ($\approx 10^{-7}$ mol dm^{-3}).

END OF MARK SCHEME · 92 marks