InkChemistry

ANSWER BOOK

Weak acids, Ka and pKa; Kw and strong bases

Acids and bases · Unit C11

AQA 3.1.12.3, 3.1.12.4 · OCR A 5.1.3

21 questions · 65 marks

NAVY EDITION

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SECTION A · Short, 1 to 2 marks

A1 WA-Q1

B1 the ionic product of water is Kw = $[H^+][OH^-]$. **B1** Kw = 1.00×10^{-14} mol² dm^{-6} at 298 K. [2]

A2 WA-Q4

M1 $[OH^-] = 0.0250$ mol dm^{-3}, so $[H^+] = Kw \div [OH^-] = 1.00 \times 10^{-14} \div 0.0250 = 4.00 \times 10^{-13}$ mol dm^{-3}. **A1** pH = $-\log(4.00 \times 10^{-13}) = 12.40$, to two decimal places. [2]

A3 WA-Q17

B1 find $[OH^-]$ from the concentration of the base, multiplying by the number of OH^- per formula unit. **B1** calculate $[H^+] = Kw \div [OH^-]$, then pH = $-\log[H^+]$. [2]

A4 WA-Q21

B1 from one reading $Ka = [H^+]^2 \div [HA]$, so $pKa = 2pH + \log[HA]$ and an error in pH is doubled in pKa. **B1** at half-neutralisation pKa equals the pH reading, so the same error appears in pKa only once. [2]

SECTION B · Standard, 3 to 5 marks

B1 WA-Q7

M1 $[H^+] = 10\text{^}(-3.10) = 7.94 \times 10^{-4}$ mol dm^{-3}. **M1** $Ka = [H^+]^2 / [HA] = (7.94 \times 10^{-4})^2 \div 0.0800$. **A1** $Ka = 7.89 \times 10^{-6}$ mol dm^{-3}, to three significant figures; allow 7.88 to 7.90. **A1** $pKa = -\log(7.89 \times 10^{-6}) = 5.10$, to two decimal places; allow error carried forward from Ka. [4]

B2 WA-Q10

B1 the dissociation $H_2O \rightleftharpoons H^+ + OH^-$ is endothermic. **B1** a rise in temperature shifts the equilibrium to the right, so Kw increases. **B1** $[H^+]$ increases, so pH = $-\log[H^+]$ falls below 7. **B1** $[OH^-]$ increases by the same amount, so $[H^+] = [OH^-]$ and the water is still neutral; reject 'neutral because pH 7'. [4]

B3 WA-Q15

B1 hydrochloric acid has the lower pH, 1.00 against 2.88, because it is fully dissociated. **B1** magnesium reacts faster with the hydrochloric acid, because its $[H^+]$ is higher. **B1** both need 25.0 cm³ of sodium hydroxide, the same volume. **B1** because as OH^- removes H^+ from ethanoic acid the equilibrium shifts right until all the acid has reacted. [4]

B4 WA-Q2

B1 $Ka = [H^+][CH_3CH_2COO^-] / [CH_3CH_2COOH]$. **B1** the units of Ka are mol dm^{-3}. **B1** pKa = $-\log$ Ka, so a larger pKa means a smaller Ka and a weaker acid. [3]

B5 WA-Q5

M1 each $Ba(OH)_2$ gives two OH^-, so $[OH^-] = 2 \times 0.0150 = 0.0300$ mol dm^{-3}. **M1** $[H^+] = 1.00 \times 10^{-14} \div 0.0300 = 3.33 \times 10^{-13}$ mol dm^{-3}. **A1** pH = 12.48, to two decimal places; 12.18 from $[OH^-] = 0.0150$ is rejected. [3]

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B6 WA-Q6 — **M1** $K_a = [H^+]^2 / [HCOOH]$, taking $[H^+] = [HCOO^-]$ and $[HCOOH] = 0.0500$ mol dm^{-3}. **M1** $[H^+] = \sqrt{(1.78 \times 10^{-4} \times 0.0500)} = 2.98 \times 10^{-3}$ mol dm^{-3}. **A1** pH = 2.53, to two decimal places. [3]

B7 WA-Q8 — **M1** $[H^+] = 10\text{^}(-3.00) = 1.00 \times 10^{-3}$ mol dm^{-3}. **M1** $[CH_3COOH] = [H^+]^2 \div K_a = (1.00 \times 10^{-3})^2 \div 1.74 \times 10^{-5}$. **A1** $[CH_3COOH] = 0.0575$ mol dm^{-3}, to three significant figures. [3]

B8 WA-Q9 — **M1** in pure water $[H^+] = \sqrt{(5.48 \times 10^{-14})} = 2.34 \times 10^{-7}$ mol dm^{-3}, giving pH 6.63. **M1** in the NaOH, $[H^+] = 5.48 \times 10^{-14} \div 0.100 = 5.48 \times 10^{-13}$ mol dm^{-3}. **A1** pH of the sodium hydroxide = 12.26, to two decimal places. [3]

B9 WA-Q11 — **B1** $[H^+] = [A^-]$, because the H^+ from the dissociation of water is negligible compared with that from the acid. **B1** $[HA]$ at equilibrium equals the initial concentration of the acid. **B1** because the acid is only slightly dissociated, so the amount of HA lost is negligible. [3]

B10 WA-Q12 — **B1** for a weak acid $[H^+] = \sqrt{(K_a \times [HA])}$, so $[H^+]$ varies with the square root of the concentration. **B1** a tenfold dilution divides $[H^+]$ by $\sqrt{10}$, about 3.2, which is a change of log $\sqrt{10} = 0.5$ in pH. **B1** dilution shifts $HA \rightleftharpoons H^+ + A^-$ to the right, so a greater fraction of the acid dissociates and $[H^+]$ falls by less than the concentration. [3]

B11 WA-Q13 — **B1** the approximation assumes $[HA]$ at equilibrium equals the initial concentration, 0.00100 mol dm^{-3}. **B1** with so large a K_a and so dilute a solution, most of the acid dissociates, about 98%. **B1** so $[HA]$ is far below 0.00100 mol dm^{-3}; the approximation gives pH 2.15, while the true pH is about 3.01. [3]

B12 WA-Q14 — **B1** chlorine is more electronegative than hydrogen and draws electron density away from the carboxylate group. **B1** the negative charge on the $ClCH_2COO^-$ ion is spread out, so the ion is more stable. **B1** the anion is therefore less ready to accept a proton back, so the equilibrium lies further to the right and K_a is larger. [3]

B13 WA-Q16 — **B1** the method assumes the acid is only slightly dissociated, so that $[HA]$ equals 1.00×10^{-4} mol dm^{-3}. **B1** at this low concentration about a third of the acid is dissociated, so $[HA]$ is well below the value used and $[H^+]$ is overestimated. **B1** the method is reliable at higher concentrations, below about 5% dissociation, but gives a pH 0.09 too low here. [3]

B14 WA-Q18 — **B1** sodium hydroxide is a strong base, fully dissociated, so $[OH^-] = 0.100$ mol dm^{-3}. **B1** ammonia is a weak base, only partially protonated in $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$, so $[OH^-]$ is much less than 0.100 mol dm^{-3}. **B1** a lower $[OH^-]$ means a higher $[H^+] = K_w \div [OH^-]$, so the pH of the ammonia is lower. [3]

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B15 WA-Q20 **M1** in pure water pH = pOH, so pH = ½pKw = 14.94 ÷ 2. **A1** the pH of pure water at 0 °C is 7.47. **A1** for the NaOH, pOH = -log(0.0100) = 2.00, so pH = 14.94 - 2.00 = 12.94. [3]

B16 WA-Q22 **M1** Ka = $[H^+][CH_3COO^-] \div [CH_3COOH]$, with $[H^+] = [CH_3COO^-]$ and $[CH_3COOH]$ taken as 0.100 mol dm^{-3}, so $[H^+]^2 = Ka \times 0.100$. **M1** $[H^+] = \sqrt{(1.74 \times 10^{-6})} = 1.32 \times 10^{-3}$ mol dm^{-3}. **A1** pH = $-\log_{10}(1.32 \times 10^{-3}) = 2.88$, to two decimal places; the two assumptions, that very little acid dissociates and that water supplies negligible H^+, should be stated. [3]

SECTION C · Long, 6 marks or more

C1 WA-Q3 **Indicative content** [6]

B1 measure the pH of a benzoic acid solution of accurately known concentration with a pH meter calibrated against buffers. **B1** convert the reading to $[H^+]$ with $[H^+] = 10^{\wedge}(-pH)$, carrying the unrounded value into the next step of the calculation. **B1** assume $[H^+] = [A^-]$, because the hydrogen ions from the dissociation of water are negligible beside those from the acid. **B1** assume $[HA]$ at equilibrium equals the concentration made up, because only a small fraction of the weak acid dissociates. **B1** calculate Ka = $[H^+]^2 \div [HA]$ in mol dm^{-3}, then pKa = -log Ka, giving pKa to two decimal places with no units. **B1** the second assumption fails if more than about 5% of the acid dissociates, as in very dilute solutions or for acids with a large Ka.

Levels

Level 3 (5–6 marks) every stage is present in order: the calibrated pH measurement on a solution of known concentration, the conversion to $[H^+]$, both assumptions stated with the reason for each, Ka and pKa calculated correctly with units handled, and the condition under which the small-dissociation assumption fails.

Level 2 (3–4 marks) the measurement, the conversion to $[H^+]$ and the Ka expression are right and one assumption is stated; why it holds, or when it fails, is left vague, or the pKa step is not finished.

Level 1 (1–2 marks) the pH is measured and Ka written down; a frequent slip is to take $[H^+]$ as equal to the acid concentration, as for a strong acid, and no assumption is justified.

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