19.1 Proton transfer: acids and bases

Name: _____ Class: _____ Date: _____ **Total: 32 marks**

KEY WORDS Brønsted–Lowry acid – • conjugate base • amphiprotic • pH pH • strong acid • weak acid

Objective

Build the skills to answer exam questions on **proton transfer** : Brønsted–Lowry acids and bases, pH, strong and weak acids, and neutralization.

You must be able to:

- identify Brønsted–Lowry acids and bases and their conjugate pairs, including amphiprotic species
- calculate pH from $[H^+]$ and use K_w
- distinguish strong and weak acids by experiment
- calculate concentrations from titration data
- use K_a, K_b, pK_a and pOH; explain buffers, salt hydrolysis and pH curves

1 Key knowledge and worked examples

Acids and bases

Brønsted–Lowry: an acid is a **proton donor**; a base is a **proton acceptor**. A **conjugate pair** differs by one H^+: CH_3COOH / CH_3COO^-. An **amphiprotic** species can donate or accept a proton (H_2O, HCO_3^-).

pH $= -\log_{10}[H^+]$; $[H^+] = 10^{-\mathrm{pH}}$. A change of 1 pH unit is a **ten-fold** change in $[H^+]$.

Water: $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 298 K. Neutral means $[H^+] = [OH^-]$.

	strong acid (HCl)	**weak acid (CH_3COOH)**
ionization	complete	partial (equilibrium)
pH at the same concentration	lower	higher
electrical conductivity	higher	lower
rate with Mg or $CaCO_3$	faster	slower

Neutralization: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. For a titration, $n = cV$ (with V in dm^3).

HL: $K_a = \dfrac{[H^+][A^-]}{[HA]}$, $pK_a = -\log K_a$; for a weak acid $[H^+] \approx \sqrt{K_a c}$. $\mathrm{pH} + \mathrm{pOH} = 14$. A **buffer** (weak acid + its conjugate base) resists pH change. At half-equivalence, $\mathrm{pH} = pK_a$.

Worked example 1: pH. $[H^+] = 0.010$ mol dm^{-3}: pH = 2.00.

Worked example 2: from K_w. $[OH^-] = 0.0010$: $[H^+] = \dfrac{1.0 \times 10^{-14}}{0.0010} = 10^{-11}$, pH = 11.00.

Answer shapes: pH to 2 decimal places; name both members of a conjugate pair.

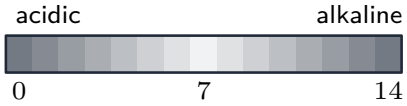
2 Practice

2.1 $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$. Identify the Brønsted–Lowry acid in the forward reaction, and give its conjugate base. [2]

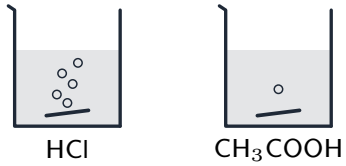
$$\mathrm{NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-}$$

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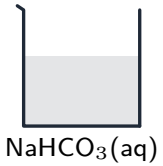
2.2 Calculate the pH of (a) a solution with $[H^+] = 0.0050$ mol dm^{-3}; (b) 0.010 mol dm^{-3} sodium hydroxide at 298 K. [3]



2.3 0.10 mol dm^{-3} hydrochloric acid has pH 1.0, but 0.10 mol dm^{-3} ethanoic acid has pH 2.9. Explain the difference, and describe one other experiment that shows which acid is weak. [3]

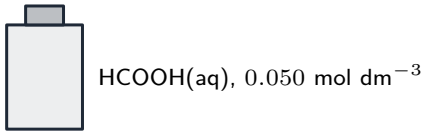


2.4 The hydrogencarbonate ion, HCO_3^-, is amphiprotic. Explain what this means, and write an equation for HCO_3^- acting as an acid with water. [3]

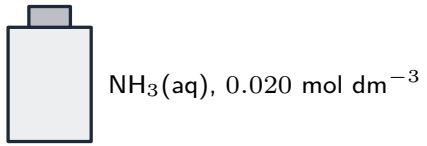


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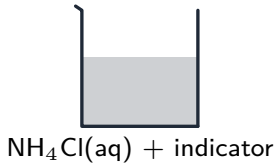
2.5 Methanoic acid has $K_a = 1.8 \times 10^{-4}$. Calculate the pH of a 0.050 mol dm^{-3} solution. [3]



2.6 Ammonia has $K_b = 1.8 \times 10^{-5}$. Calculate the pH of a 0.020 mol dm^{-3} solution at 298 K. [3]

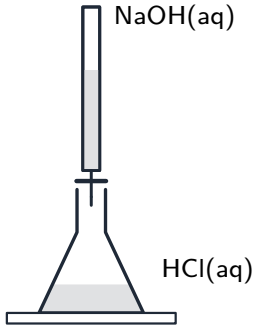


2.7 Predict whether a solution of ammonium chloride is acidic, neutral or alkaline. Explain with an equation. [3]



3 Exam-style questions

3.1 A student titrates 25.0 cm^3 of hydrochloric acid with 0.100 mol dm^{-3} sodium hydroxide. The mean titre is 21.50 cm^3.

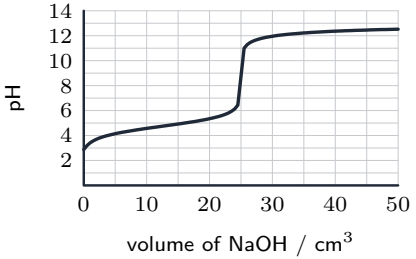


(a) Write the ionic equation for the neutralization. [1]

(b) Calculate the concentration of the hydrochloric acid. [3]

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3.2 25.0 cm^3 of 0.100 mol dm^{-3} ethanoic acid is titrated with 0.100 mol dm^{-3} sodium hydroxide. The pH curve is shown.



(a) Use the curve to find pK_a of ethanoic acid, and calculate K_a. [2]

(b) Explain why the pH at the equivalence point is above 7. [2]

(c) Phenolphthalein changes colour between pH 8.2 and 10.0; methyl orange between 3.1 and 4.4. Choose the better indicator and explain. [2]

(d) Explain how the mixture at 12.5 cm^3 resists a change in pH when a little acid is added. [2]

Solutions

Marking: **M1** a correct method, **A1** the correct answer, **B1** an independent mark.

Each answer shows the steps, then the **result**.

2.1

H_2O; OH^-

2.2

(a) **2.30**; (b) $[H^+] = \dfrac{1.0 \times 10^{-14}}{0.010} = 1.0 \times 10^{-12}$; pH = **12.00**

2.3

HCl ionizes completely, ethanoic acid only partly; so ethanoic acid has a lower $[H^+]$; ethanoic acid conducts electricity less well / reacts more slowly with Mg (fewer bubbles)

2.4

it can donate a proton; and it can accept a proton; $HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+$

2.5

$[H^+] = \sqrt{K_a c} = \sqrt{1.8 \times 10^{-4} \times 0.050} = 3.0 \times 10^{-3}$; pH = **2.52**

2.6

$[OH^-] = \sqrt{1.8 \times 10^{-5} \times 0.020} = 6.0 \times 10^{-4}$; pOH = 3.22; pH = **10.78**

2.7

acidic; NH_4^+ is a weak acid / the conjugate acid of a weak base; $NH_4^+ \rightleftharpoons NH_3 + H^+$

3.1

(a) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(b) $n(NaOH) = 0.100 \times 0.02150 = 2.150 \times 10^{-3}$ mol; $n(HCl) = 2.150 \times 10^{-3}$ mol; $c = \dfrac{2.150 \times 10^{-3}}{0.0250} =$ **0.0860 mol dm^{-3}**

3.2

(a) at half-equivalence (12.5 cm^3) pH = pK_a = 4.74 (accept 4.7–4.8); $K_a = 10^{-4.74} =$ **1.8×10^{-5}**

(b) the solution contains ethanoate ions, a weak base; $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

(c) phenolphthalein; its range lies within the steep part of the curve

(d) it is a buffer with equal amounts of CH_3COOH and CH_3COO^-; added H^+ reacts with CH_3COO^- to form CH_3COOH, so $[H^+]$ hardly changes

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