18.1 Proton transfer: acids and bases

Name: _____ Class: _____ Date: _____ **Total: 24 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

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

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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).

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.

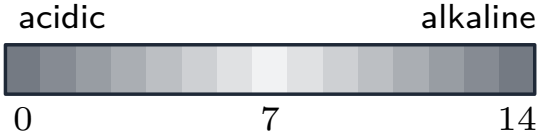
Answer shapes: pH to 2 decimal places; name both members of a conjugate pair.

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]

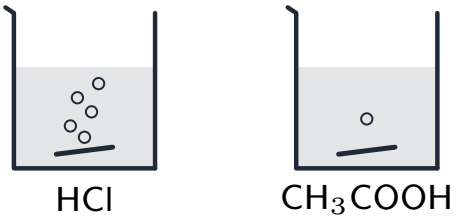
$$\mathrm{NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-}$$

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]

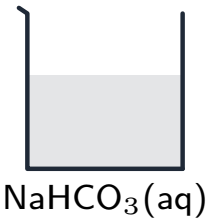


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2.3 0.10 mol dm^{-3} hydrochloric acid has pH 1.0, but 0.10 mol dm^{-3} ethanoic acid has pH 2.9. Explain the difference, and describe one other experiment that shows which acid is weak. [3]



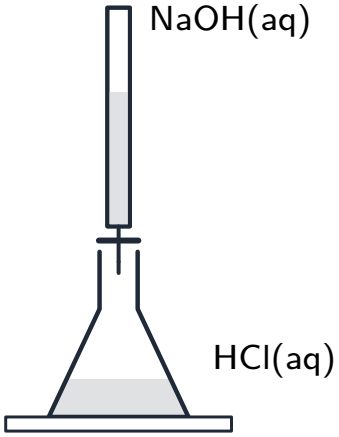
2.4 The hydrogencarbonate ion, HCO_3^-, is amphiprotic. Explain what this means, and write an equation for HCO_3^- acting as an acid with water. [3]



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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]

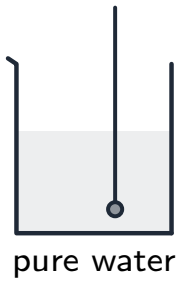
(b) Calculate the concentration of the hydrochloric acid. [3]

(c) Calculate the pH of the hydrochloric acid. [1]

(d) State the pH of the solution at the equivalence point at 298 K. [1]

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3.2 The ionic product of water, K_w, is 1.0×10^{-14} at 298 K and 5.5×10^{-14} at 323 K.



(a) Write the expression for K_w. [1]

(b) Calculate the pH of pure water at 323 K. [2]

(c) Explain why pure water at 323 K is still neutral. [1]

(d) Deduce whether the ionization of water is exothermic or endothermic. Explain your answer. [2]

(e) State how many times greater $[H^+]$ is in a solution of pH 4 than in a solution of pH 6. [1]

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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; $HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+$

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; $c = \dfrac{2.150 \times 10^{-3}}{0.0250} = \mathbf{0.0860\ mol\ dm^{-3}}$

(c) $-\log 0.0860 = \mathbf{1.07}$

(d) 7

3.2

(a) $K_w = [H^+][OH^-]$

(b) $[H^+] = \sqrt{5.5 \times 10^{-14}} = 2.3 \times 10^{-7}$; pH = **6.63**

(c) $[H^+] = [OH^-]$

(d) endothermic; K_w increases with temperature, so the equilibrium shifts right when heated

(e) 100

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