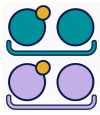
Acids, Bases and pH Map

Cambridge International AS & A Level Chemistry · Topic 25 Equilibria · 25.1 Acids, bases and pH

1 Conjugate pairs

25.1.1, 25.1.2

- **Acid**: proton (H^+) donor
- Differ by **one** H^+
- Acid → conjugate base
- Base → conjugate acid
- Pairs sit **across** the arrow



2 pH

25.1.3

- **$pH = -\log_{10}[H^+]$**
- **$[H^+] = 10^{-pH}$**
- pH to **2 d.p.**, no unit
- One unit = **×10** in $[H^+]$
- Can be below 0

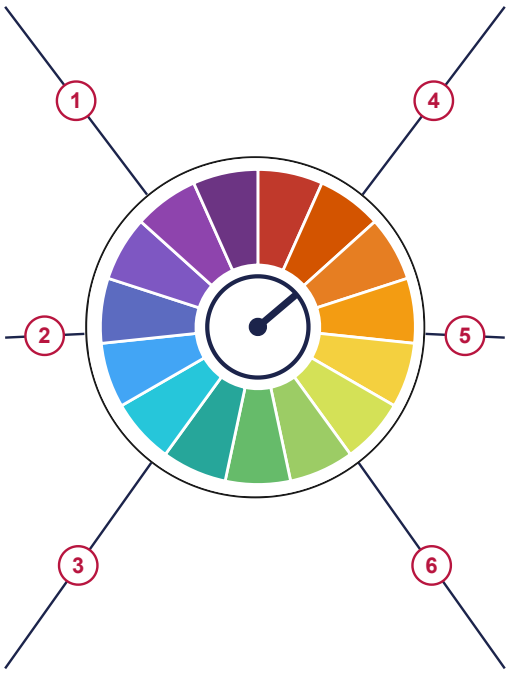


3 Strong acids

25.1.4

- **Fully dissociated**
- HCl, HNO_3: $[H^+]$ = [acid]
- pH = $-\log_{10}$[acid]
- Dilute ×10: pH **+1**
- Strong ≠ concentrated

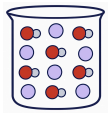




4 K_w and strong alkalis

25.1.3, 25.1.4

- **$K_w = [H^+][OH^-]$**
- 1.00×10^{-14} mol² dm⁻⁶ (298 K)
- $Ba(OH)_2$: **two** OH^-
- **$[H^+] = K_w \div [OH^-]$**
- Heating raises K_w



5 K_a and pK_a

25.1.3

- **Partially dissociated**: $HA \rightleftharpoons H^+ + A^-$
- **$K_a = [H^+][A^-] \div [HA]$**
- Units: **mol dm⁻³**
- **$pK_a = -\log_{10}K_a$**
- Smaller pK_a: **stronger**



6 pH of a weak acid

25.1.4

- **$[H^+] = [A^-]$** (water ignored)
- **[HA] ≈ c** (little dissociates)
- **$[H^+] = \sqrt{(K_a \times c)}$**
- Backwards: $c = [H^+]^2 \div K_a$
- Check: under about 5 %



pH at 298 K

strong: $[H^+]$ from the acid or from $K_w = 1.00 \times 10^{-14}$; weak: $[H^+] = \sqrt{(K_a c)}$

Solution	Data	Working	pH
0.0250 mol dm⁻³ HNO_3	strong	$[H^+] = 0.0250$	1.60
0.0500 mol dm⁻³ NaOH	strong	$[H^+] = K_w \div 0.0500 = 2.00 \times 10^{-13}$	12.70
0.0150 mol dm⁻³ $Ba(OH)_2$	strong	$[OH^-] = 0.0300$; $[H^+] = 3.33 \times 10^{-13}$	12.48
0.100 mol dm⁻³ methanoic acid	pK_a 3.75	$[H^+] = \sqrt{(1.78 \times 10^{-4} \times 0.100)}$	2.38
0.100 mol dm⁻³ ethanoic acid	pK_a 4.76	$[H^+] = \sqrt{(1.74 \times 10^{-5} \times 0.100)}$	2.88

Worked examples

Find the pairs

- $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$
- HNO_3 / NO_3^-
- **H_3O^+ / H_2O**

Neutral when hot

- 100 °C: $K_w = 5.13 \times 10^{-13}$
- pH of pure water 6.14
- **$[H^+] = [OH^-]$: neutral**

K_a from pH

- 0.100 mol dm⁻³ HA, pH 2.88
- $[H^+] = 1.32 \times 10^{-3}$
- **$K_a = 1.74 \times 10^{-5}$ mol dm⁻³**

Buffer solutions follow in 25.1; then the solubility product, Ksp, and the common ion effect.

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