Topic 12: Acid-Base Equilibria

Edexcel A Level Chemistry • Paper 1 (9CH0/01)

12.1 Know that a Brønsted–Lowry acid is a proton donor and a Brønsted–Lowry base is a proton acceptor.

Brønsted–Lowry acid: a species that donates a proton (H^+) to another species.

Brønsted–Lowry base: a species that accepts a proton (H^+) from another species.

This definition is broader than the Arrhenius definition and includes species like NH_3 (base, accepts H^+) that do not contain OH^-.

12.2 Know that acid-base reactions involve the transfer of protons.

In every Brønsted–Lowry acid-base reaction, a proton (H^+) is transferred from the acid to the base. E.g. $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$: HCl donates H^+ (acid), NH_3 accepts H^+ (base).

12.3 Be able to identify Brønsted–Lowry conjugate acid-base pairs.

A conjugate pair differs by one proton.

The conjugate base of an acid is formed when the acid loses H^+ (e.g. HCl/Cl^-).

The conjugate acid of a base is formed when the base gains H^+ (e.g. NH_3/NH_4^+).

In $HA + B \quad A^- + BH^+$: conjugate pairs are HA/A^- and B/BH^+.

12.4 Be able to define the term 'pH'.

$pH = -\log_{10}[H^+]$

pH is a logarithmic scale that measures the concentration of hydrogen ions in solution. A decrease of 1 pH unit = a 10-fold increase in $[H^+]$. Lower pH = more acidic; higher pH = more alkaline.

12.5 Be able to calculate pH from hydrogen ion concentration.

$pH = -\log_{10}[H^+]$

E.g. If $[H^+] = 0.01$ mol dm^{-3}, $pH = -\log(0.01) = 2.00$.

Always give pH to 2 decimal places in calculations.

12.6 Be able to calculate the concentration of hydrogen ions from pH, using $[H^+] = 10^{-p}$.

$[H^+] = 10^{-p}$

E.g. If pH = 3.40, then $[H^+] = 10^{-3.40} = 3.98 \times 10^{-4}$ mol dm^{-3}.

This is the inverse of the pH calculation.

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12.7 Understand the difference between a strong acid and a weak acid in terms of degree of dissociation.

A strong acid is fully (completely) dissociated in aqueous solution – every molecule releases H^+. E.g. HCl → $H^+ + Cl^-$ (use a single arrow →).

A weak acid is partially dissociated in aqueous solution – an equilibrium is established. E.g. $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ (use an equilibrium arrow ⇌). The degree of dissociation is small.

12.8 Be able to calculate the pH of a strong acid.

For a strong monobasic acid (e.g. HCl): $[H^+]$ = concentration of the acid (fully dissociated). pH = $-\log_{10}[H^+]$.

For a strong dibasic acid (e.g. H_2SO_4): $[H^+]$ = 2 × concentration of the acid (each molecule releases $2H^+$).

12.9 Be able to deduce the expression for the acid dissociation constant, Ka, for a weak acid and carry out relevant calculations.

For the weak acid $HA \rightleftharpoons H^+ + A^-$:

$K_a = [H^+][A^-] / [HA]$

Ka is the acid dissociation constant – it measures the strength of a weak acid. A larger Ka means a stronger weak acid (more dissociated). Units are usually mol dm^{-3}.

12.10 Be able to calculate the pH of a weak acid making relevant assumptions.

Assumptions for weak acid pH calculation:

- $[H^+] = [A^-]$ (each HA that dissociates produces one of each).
- [HA] at equilibrium ≈ initial [HA] (dissociation is so small that the change is negligible).

Method: $K_a = [H^+]^2 / [HA]$, so $[H^+] = \sqrt{(K_a \times [HA])}$, then pH = $-\log[H^+]$.

State both assumptions in exam answers for full marks.

12.11 Be able to define the ionic product of water, Kw.

Water undergoes autoionisation: $H_2O \rightleftharpoons H^+ + OH^-$

$K_w = [H^+][OH^-]$

At 25°C, $K_w = 1.00 \times 10^{-14}$ mol^2 dm^{-6}.

Kw increases with temperature because the forward reaction is endothermic.

12.12 Be able to calculate the pH of a strong base from its concentration, using Kw.

For a strong base like NaOH: $[OH^-]$ = concentration of base (fully dissociated).

$[H^+] = K_w / [OH^-]$

Then: pH = $-\log_{10}[H^+]$

E.g. 0.10 mol dm^{-3} NaOH: $[H^+] = 1.00 \times 10^{-14} / 0.10 = 1.00 \times 10^{-13}$, pH = 13.00.

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12.13 Be able to define the terms 'pKa' and 'pKw'.

$pKa = -\log_{10}(Ka)$ and $Ka = 10^{-pKa}$

$pKw = -\log_{10}(Kw)$ and at 25°C, pKw = 14.00.

A lower pKa means a stronger weak acid (larger Ka). The p-notation is a convenient way to handle very small numbers.

12.14 Be able to analyse data from experiments: (i) measuring pH of strong/weak acids, bases, and salts, (ii) comparing pH of strong and weak acids after dilution.

(i) A strong acid (e.g. 0.1 M HCl) has a lower pH than a weak acid of the same concentration (e.g. 0.1 M CH_3COOH) because it is fully dissociated. Strong bases have higher pH than weak bases of the same concentration.

(ii) Diluting a strong acid by a factor of 10 increases pH by exactly 1 unit (e.g. pH 1 → pH 2). Diluting a weak acid by a factor of 10 increases pH by approximately 0.5 units because the equilibrium shifts to produce more H^+, partially compensating for the dilution.

12.15 Be able to calculate Ka for a weak acid from experimental data given the pH of a solution containing a known mass of acid.

Method: (1) Calculate $[H^+]$ from pH: $[H^+] = 10^{-pH}$. (2) $[A^-] = [H^+]$ (from dissociation). (3) Calculate initial [HA] from mass, Mr, and volume. (4) [HA] at equilibrium = initial [HA] – $[H^+]$ (or assume ≈ initial [HA] if $[H^+]$ is very small compared to [HA]). (5) Substitute into $Ka = [H^+][A^-] / [HA]$.

12.16 Be able to draw and interpret titration curves using all combinations of strong and weak monobasic acids and bases.

Key features of titration curves:

- Strong acid + strong base: sharp vertical section at equivalence point; equivalence pH = 7.
- Weak acid + strong base: starts at higher pH; gradual rise then sharp at equivalence; equivalence pH > 7 (due to hydrolysis of conjugate base); buffer region before equivalence point.
- Strong acid + weak base: equivalence pH < 7.
- Weak acid + weak base: no sharp change at equivalence; very difficult to determine end point.

The equivalence point is where stoichiometric amounts have reacted.

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12.17

Be able to select a suitable indicator, using a titration curve and appropriate data.

A suitable indicator must have its colour change (pKa range) within the vertical (steep) section of the titration curve.

- Strong acid/strong base: phenolphthalein or methyl orange (sharp change covers both ranges).
- Weak acid/strong base: phenolphthalein only (equivalence pH > 7; methyl orange changes too early).
- Strong acid/weak base: methyl orange only (equivalence pH < 7; phenolphthalein changes too late).
- Weak/weak: no suitable indicator (no sharp change).

12.18

Know what is meant by the term 'buffer solution'.

A buffer solution is one that resists changes in pH when small amounts of acid or base are added, or when the solution is diluted. Buffer solutions are made from:

- A weak acid and its conjugate base (e.g. CH_3COOH + CH_3COONa) – acidic buffer.
- A weak base and its conjugate acid (e.g. NH_3 + NH_4Cl) – basic buffer.

12.19

Understand the action of a buffer solution.

Acidic buffer (e.g. CH_3COOH / CH_3COO^-):

When H^+ is added: the conjugate base (CH_3COO^-) reacts with the added H^+ to form the weak acid: $CH_3COO^- + H^+ \rightarrow CH_3COOH$. This removes the added H^+, so pH barely changes.

When OH^- is added: the weak acid (CH_3COOH) reacts with the added OH^-: $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$. This removes the added OH^-.

The large reservoir of weak acid and conjugate base allows the buffer to absorb small additions of either acid or base.

12.20

Be able to calculate the pH of a buffer solution given appropriate data.

Method: Use $K_a = [H^+][A^-] / [HA]$.

Rearrange: $[H^+] = K_a \times [HA] / [A^-]$

Then: $pH = -\log[H^+]$.

[HA] = concentration of the weak acid in the buffer. $[A^-]$ = concentration of the conjugate base (salt) in the buffer. Since both are in the same solution, you can use moles instead of concentrations (the volumes cancel).

Alternatively: $pH = pK_a + \log([A^-] / [HA])$ (Henderson–Hasselbalch equation).

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12.21 Be able to calculate the concentrations of solutions required to prepare a buffer solution of a given pH.

Method: Rearrange $[H^+] = Ka \times [HA] / [A^-]$ to find the required ratio of [HA] to $[A^-]$.

(1) Calculate $[H^+]$ from the target pH. (2) Ratio: $[HA]/[A^-] = [H^+]/Ka$. (3) Choose convenient concentrations/volumes that give this ratio. You only need the ratio to set the pH; the actual concentrations determine the capacity of the buffer.

12.22 Understand how to use a weak acid–strong base titration curve to: (i) demonstrate buffer action, (ii) determine Ka from the pH at the half-equivalence point.

(i) Buffer region: the relatively flat section of the curve before the equivalence point shows buffer action – pH changes only gradually as base is added because the solution contains both the weak acid and its conjugate base.

(ii) At the half-equivalence point (half the acid neutralised): $[HA] = [A^-]$, so $Ka = [H^+]$, meaning pH = pKa. Read the pH from the curve at exactly half the equivalence volume to determine Ka.

12.23 Understand why there is a difference in enthalpy changes of neutralisation values for strong and weak acids.

Strong acid + strong base: $\Delta H^n_{eut} \approx -57.1$ kJ mol^{-1} (the standard value for $H^+ + OH^- \rightarrow H_2O$).

Weak acid + strong base: ΔH^n_{eut} is less exothermic (less negative) because some energy is used to dissociate the weak acid further (the dissociation of the weak acid is endothermic). This endothermic dissociation partially offsets the exothermic neutralisation.

12.24 Understand the roles of carbonic acid molecules and hydrogencarbonate ions in controlling the pH of blood.

Blood pH is maintained at approximately 7.4 by the carbonate buffer system:

$CO_2(aq) + H_2O(l) \quad H_2CO_3(aq) \quad H^+(aq) + HCO_3^-(aq)$

If blood pH falls (too much H^+): HCO_3^- ions react with excess $H^+ \rightarrow H_2CO_3 \rightarrow CO_2$ (exhaled). This removes H^+ and raises pH.

If blood pH rises (too much OH^- or too little H^+): H_2CO_3 dissociates to release H^+, lowering the pH.

The equilibrium between CO_2, H_2CO_3 and HCO_3^- acts as the buffer. Breathing rate also helps regulate CO_2 levels.

Finding the Ka Value for a Weak Acid

ABOUT THESE NOTES

Notes written by Chloe Archard (MSc Biochemistry, Oxford-trained chemist, 20+ years teaching experience at top UK independent schools). For 1-1 tuition, revision courses, and more free resources: chemistrywithchloe.com

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