Topic 8.4: Acid-Base Reactions and Buffers

Revision Notes · AP Chemistry

UNIT 8 · ACIDS AND BASES

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Neutralization: One Pattern, Three Pairings

When an acid meets a base, a proton changes hands, and the transfer runs to completion wherever a strong partner is involved. What remains is a salt, usually with water, and the pairing decides its character.

Strong acid with strong base:

$HCl + NaOH \rightarrow NaCl + H_2O$

Neither ion left behind reacts with water, so the solution is neutral.

Weak acid with strong base:

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

The acetate ion is a weak base ($K_b = 5.6 \times 10^{-10}$), so the solution is basic.

Strong acid with weak base:

$HCl + NH_3 \rightarrow NH_4Cl$

The ammonium ion is a weak acid ($K_a = 5.6 \times 10^{-10}$), so the solution is acidic.

Net Ionic Equations: Strip the Spectators

A net ionic equation lists only the species that change. Strong acids, strong bases and soluble salts are written as separate ions; weak acids, weak bases and water stay whole, as they exist in solution.

For $HCl + NaOH$, Na^+ and Cl^- appear unchanged on both sides and cancel:

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$H_3O^+ + OH^- \rightarrow 2H_2O$

This is often shortened to:

$H^+ + OH^- \rightarrow H_2O$

For CH_3COOH + NaOH the acetic acid must not be split:

$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

Splitting it would claim full ionization, which its Ka of 1.8×10^{-5} denies.

For HCl + NH_3 the weak base stays whole:

$H_3O^+ + NH_3 \rightarrow NH_4^+ + H_2O$

The reverse transfer is what strong base does to an ammonium salt:

$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$

What a Buffer Is

A buffer is a solution that resists pH change when small amounts of strong acid or strong base are added. It needs appreciable amounts of a weak acid with its conjugate base, HA with A^-, or a weak base with its conjugate acid, B with BH^+.

Both members must be present in quantity: CH_3COOH with CH_3COONa, NH_3 with NH_4Cl, and H_2CO_3 with $NaHCO_3$ all qualify.

HCl with NaCl does not. Cl^-, the conjugate base of a strong acid, has no proton appetite: it never takes H_3O^+ back, so an attack in either direction meets no partner.

How a Buffer Absorbs an Attack

Attack by acid. Added H_3O^+ is consumed by the base member:

$CH_3COO^- + H_3O^+ \rightarrow CH_3COOH + H_2O$

The strong acid becomes the weak acid already present, so free H_3O^+ barely accumulates.

Attack by base. Added OH^- is consumed by the acid member:

$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

The strong base ends as water plus more conjugate base.

In both directions the members trade places, so the ratio and the pH barely move.

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The resistance has a hard limit: an attack is absorbed only while both members survive. Add more H_3O^+ than there is A^-, or more OH^- than there is HA, and the excess acts alone.

Buffer pH: The Ratio Runs the Number

A buffer is the Topic 8.3 equilibrium, both sides loaded:

$Ka = [H_3O^+][A^-] / [HA]$

Rearranging gives:

$[H_3O^+] = Ka \times [HA] / [A^-]$

Only the ratio matters: double both members and $[H_3O^+]$ is unchanged.

In log form the same statement reads:

$pH = pKa + \log([A^-] / [HA])$

When the two members are present in equal amounts the log term is zero, so the pH equals the pKa. An equimolar acetate buffer sits at pH 4.74.

Choosing the pair. A buffer holds its pH well only near its acid's pKa. For pH 7.0, $H_2PO_4^-$ / HPO_4^{2-} (pKa 7.21) is the pick; CH_3COOH / CH_3COO^- (pKa 4.74) would need a ratio so lopsided that one member nearly vanishes. The closest pKa wins.

A buffer by half-neutralization. Mixing a weak acid with half the strong base needed to neutralize it converts half of the HA into A^-, leaving equal amounts, so the pH equals the pKa. A half-neutralized weak acid is a ready-made buffer.

Stoichiometry First: A Buffer After an Addition

Adding strong acid or base to a buffer is a mole problem first, an equilibrium problem second: let the added ion react completely with the member it attacks, recount HA and A^-, then put the new amounts into the ratio. Work in moles: both members share one volume, so it cancels. Confirm both members survive: a negative amount means the leftover strong acid or base sets the pH alone.

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Worked Examples

Example 1: An acetate buffer away from the midpoint. A buffer holds 0.25 mol L^{-1} CH_3COOH and 0.15 mol L^{-1} CH_3COO^-, with Ka = 1.8×10^{-5}. The ratio route gives $[H_3O^+] = 1.8 \times 10^{-5} \times 0.25 / 0.15 = 3.0 \times 10^{-5}$ mol L^{-1}, so pH = 4.52. The full equilibrium returns the same value: ionization shifts the ratio by only about 0.02 percent. The acid member outweighs the base member, so the pH sits below the pKa.

Example 2: An ammonia buffer under acid attack. One litre of buffer holds 0.20 mol NH_3 and 0.20 mol NH_4Cl, with Kb of NH_3 = 1.8×10^{-5}. Equal members put the pH at 14.00 − pKb = 9.26. Now 0.030 mol of HCl is added. Its H_3O^+ converts NH_3 into NH_4^+ mole for mole: NH_3 falls to 0.17 mol and NH_4^+ rises to 0.23 mol. Then $[OH^-] = Kb \times [NH_3] / [NH_4^+] = 1.8 \times 10^{-5} \times 0.17 / 0.23 = 1.3 \times 10^{-5}$ mol L^{-1}, giving pOH 4.88 and pH 9.12, a fall of 0.14. The same 0.030 mol of HCl in a litre of plain water gives pH 1.52. The buffer absorbed the attack almost silently.

Example 3: An HCN buffer. A buffer holds 0.15 mol L^{-1} HCN and 0.10 mol L^{-1} NaCN, with Ka of HCN = 4.9×10^{-10}, so pKa = 9.31. The ratio gives pH = 9.31 + log(0.10 / 0.15) = 9.31 − 0.18 = 9.13, below the pKa because the base member is scarcer.

Example 4: An acetate buffer under base attack. One litre of buffer holds 0.10 mol CH_3COOH and 0.10 mol CH_3COO^-. Adding 0.010 mol of NaOH converts HA into A^- mole for mole: HA falls to 0.090 mol and A^- rises to 0.11 mol. Then pH = 4.74 + log(0.11 / 0.090) = 4.74 + 0.09 = 4.83, a rise of 0.09.

Traps That Cost Marks in Topic 8.4

"HCl and NaCl are conjugates, so together they buffer." They are conjugates, and it does not help: Cl^- never takes a proton back. The pair must be weak.

Splitting the weak acid in the net ionic equation. Writing CH_3COOH as $H^+ + CH_3COO^-$ before reacting it with OH^- is the strong-acid reflex; the whole molecule is the reacting species.

Reaching for an ICE table right after an addition. An addition changes the moles of both members. Recount first, then apply the ratio; an equilibrium calculation on the original amounts answers a question nobody asked.

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"The buffer kept the pH exactly the same." It never does: Example 2 moved 0.14 units where an unbuffered litre moved about five and a half.

Reading a salt by its name instead of its ions. "A salt is neutral" fails twice here: NH_4Cl is acidic because NH_4^+ donates protons to water, and CH_3COONa is basic because CH_3COO^- accepts them.

Turning the ratio upside down. $[H_3O^+] = Ka \times [HA] / [A^-]$: more acid member means more H_3O^+ and a lower pH. The free check is the midpoint: equal members must return pH = pKa.

Key Relationships to Remember

- Strong acid + strong base gives a neutral salt and water; weak acid + strong base leaves a basic conjugate base; strong acid + weak base leaves an acidic conjugate acid
- Net ionic rule: split strong electrolytes only, keep weak acids and bases whole, cancel ions unchanged on both sides
- A buffer is appreciable amounts of a weak acid with its conjugate base, or a weak base with its conjugate acid
- Added H_3O^+ is consumed by A^-; added OH^- is consumed by HA
- $[H_3O^+] = Ka \times [HA] / [A^-]$, equivalently $pH = pKa + \log([A^-] / [HA])$; equal members give pH = pKa
- Choose the pair whose pKa lies closest to the target pH
- After any addition: stoichiometry first, ratio second; an exhausted member hands the pH to the leftover strong acid or base

Continue Revising

Topic 8.4 Quiz: Acid-Base Reactions and Buffers

Topic 8.4 Worksheet

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