Topic 4.8: Introduction to Acid-Base Reactions

Revision Notes · AP Chemistry

UNIT 4 · CHEMICAL REACTIONS

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Two Definitions of an Acid

The older definition is Arrhenius: an acid produces H^+ ions in water, a base produces OH^- ions in water. Sodium hydroxide is an Arrhenius base because it releases OH^- directly. But ammonia, NH_3, contains no OH^- to release, yet behaves as a base: the Arrhenius picture has a gap.

The Brønsted-Lowry definition, the CED's, is built on proton transfer: an acid is a proton (H^+) donor, a base a proton acceptor. A proton can only be donated if something accepts it, so acid and base always appear together.

In Water, the Proton Travels as H_3O^+

A bare H^+ never exists alone in water: it attaches to a water molecule, forming hydronium, H_3O^+. $H^+(aq)$ is shorthand for $H_3O^+(aq)$.

Hydrogen chloride dissolving: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$. HCl donates the proton (the acid); water accepts it (the base).

With ammonia, water's role reverses: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$, water donating the proton and acting as the acid. The OH^- was made by the transfer, not released from the ammonia.

No acid stronger than H_3O^+ survives in water: it hands its proton to water instead, making H_3O^+ the strongest acid present in aqueous solution. In $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ the equilibrium therefore lies well to the left: only a small fraction of NH_4^+ ions have donated at any moment.

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Conjugate Pairs Differ by Exactly One Proton

When an acid donates its proton, what remains is its conjugate base. When a base accepts a proton, what forms is its conjugate acid. A conjugate pair is two species differing by exactly one H^+, their charges differing by one too.

Conjugate base: remove one H^+ (H_2SO_4 gives HSO_4^-, not SO_4^{2-}). Conjugate acid: add one H^+ (NH_3 gives NH_4^+). Every proton transfer contains two pairs: in $HCl + H_2O \rightarrow H_3O^+ + Cl^-$, HCl/Cl^- and H_3O^+/H_2O.

A weak acid together with its conjugate base forms a buffer: the acid neutralises added OH^-, the conjugate base added H^+, so the pH changes far less than in water alone.

Amphiprotic Species Play Both Roles

An amphiprotic species donates a proton in one reaction, accepts one in another; its partner decides which. Water accepts from HCl, donates to NH_3.

Hydrogencarbonate, HCO_3^-, is amphiprotic: with water it acts as a base, $HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$; with hydroxide as an acid, $HCO_3^- + OH^- \rightarrow CO_3^{2-} + H_2O$. HSO_4^- and $H_2PO_4^-$ behave the same way.

Strong Means Fully Ionised, Weak Means Barely

A strong acid donates its proton completely: no intact acid molecules remain. The strong acids are HCl, HBr, HI, HNO_3, $HClO_4$ and the first proton of H_2SO_4; single arrow. Strong bases ionise completely too: Group 1 hydroxides, such as NaOH and KOH, and $Ba(OH)_2$ are entirely cations and OH^- ions.

A weak acid ionises only slightly: $CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$ is an equilibrium, and in 0.10 mol/L solution only about one molecule in a hundred has ionised at any moment. A weak base accepts protons only partially: most NH_3 in solution is still NH_3.

Strength is the fraction ionised; concentration is the amount dissolved. Independent ideas: a concentrated weak acid can hold more H_3O^+ than a very dilute strong acid.

Neutralisation: A Proton Meets Hydroxide

An acid reacting with an OH^- base produces a salt and water: $HCl + NaOH \rightarrow NaCl + H_2O$, the salt pairing the base's cation with the acid's anion.

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For any strong acid with any strong base the net ionic equation is Topic 4.2's: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$; the spectators differ, the proton transfer never does. With hydronium: $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$.

If the acid is weak it stays whole in the net ionic equation, because most of it is whole in the flask: $CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)$. If the base is ammonia, no OH^- appears: $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$; with hydrochloric acid the only product is the salt NH_4Cl. Neutralisation is the proton transfer, not the appearance of water.

Strength changes how ionised an acid is, not its total proton capacity: a mole of CH_3COOH neutralises exactly as much NaOH as a mole of HCl.

Polyprotic Acids Donate One Proton at a Time

A polyprotic acid has more than one donatable proton: H_2SO_4 has two, H_3PO_4 three. They leave one at a time, in separate steps, never together.

Sulfuric acid's first step is complete, $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$; the second is an equilibrium, $HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$. Hence the conjugate base of H_2SO_4 is HSO_4^-: one proton removed, one step taken.

Carbonate runs the same staircase in reverse: $CO_3^{2-} + H^+ \rightarrow HCO_3^-$, then $HCO_3^- + H^+ \rightarrow H_2O + CO_2$ (through H_2CO_3, which falls apart). Together: $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$, the net ionic equation behind the Topic 4.7 result.

Worked Examples

Example 1: Labelling a proton transfer. In $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$, NH_3 accepts a proton: it is the base, and NH_4^+ is its conjugate acid. H_2O donates the proton: it is the acid, and OH^- its conjugate base. Two pairs, each differing by one proton.

Example 2: Three neutralisations, three net ionic equations. Nitric acid with potassium hydroxide: K^+ and NO_3^- spectate: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. Acetic acid with sodium hydroxide: Na^+ spectates, the weak acid stays whole: $CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)$. Hydrochloric acid with ammonia: Cl^- spectates, the proton lands on ammonia: $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$. One family, three net equations, because what is present in each flask differs.

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Example 3: Carbonate, one proton at a time. A solution contains 0.0200 mol of CO_3^{2-}. Adding 0.0200 mol of H^+ runs the first step only, $CO_3^{2-} + H^+ \rightarrow HCO_3^-$: 0.0200 mol of HCO_3^-, no gas. A second 0.0200 mol of H^+ runs the second step, $HCO_3^- + H^+ \rightarrow H_2O + CO_2$, and the 0.0200 mol of CO_2 escapes: 0.0200 × 22.4 = 0.448 L at STP. Total acid: 0.0400 mol, two protons per carbonate.

Traps That Cost Marks in Topic 4.8

Removing both protons at once. The conjugate base of H_2SO_4 is HSO_4^-. A conjugate pair differs by exactly one proton, so SO_4^{2-}, two protons away, cannot be the answer.

Miscounting the gap in a pair. H_2CO_3 and CO_3^{2-} are not a conjugate pair: two protons apart. The real pairs are H_2CO_3/HCO_3^- and HCO_3^-/CO_3^{2-}. H_2O and OH^-, by contrast, differ by one proton and are a genuine pair.

Splitting weak acids and bases in ionic equations. CH_3COOH and NH_3 are written whole. Only strong acids and strong bases appear as separated ions, exactly as in Topic 4.2.

Confusing strong with concentrated. Strong describes how completely a substance ionises; concentrated, how much is dissolved. Dilute HCl is still strong; concentrated CH_3COOH is still weak.

Expecting water, or OH^-, in every neutralisation. $NH_3 + HCl \rightarrow NH_4Cl$ forms a salt, no water, and its net ionic equation has no OH^-: the base took the proton itself.

Pinning a permanent label on a substance. Water is the base when HCl dissolves in it, the acid when NH_3 does; HCO_3^- is a base toward water, an acid toward OH^-. Assign roles from the equation in front of you.

Forgetting the second water. $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$: written with hydronium, neutralisation makes two waters, not one.

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Key Relationships to Remember

- Brønsted-Lowry: an acid donates H^+, a base accepts H^+; every acid-base reaction is a proton transfer needing both partners. Arrhenius is the water-only version: acid gives H^+, base gives OH^-
- In water the proton travels as H_3O^+ ($H^+(aq)$ is shorthand), the strongest acid that can exist in aqueous solution
- A conjugate pair differs by exactly one H^+: base + H^+ gives its conjugate acid, acid − H^+ gives its conjugate base, the charge shifting by one each way
- Every proton transfer contains two conjugate pairs; a weak acid with its conjugate base forms a buffer resisting pH change
- Strong acids and bases ionise completely and are written as ions; weak ones reach equilibrium and stay mostly intact; strength (fraction ionised) is not concentration (amount dissolved)
- Neutralisation net ionic equations: $H^+ + OH^- \rightarrow H_2O$ (strong with strong); $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$ (weak acid); $NH_3 + H^+ \rightarrow NH_4^+$ (ammonia)
- Polyprotic acids donate stepwise, $H_2SO_4 \rightarrow HSO_4^- \rightarrow SO_4^{2-}$; carbonate accepts stepwise, $CO_3^{2-} \rightarrow HCO_3^- \rightarrow H_2O + CO_2$, so acid plus carbonate gives a salt, water and carbon dioxide

Continue Revising

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