Topic 8.2: pH and pOH of Strong Acids and Bases

Revision Notes · AP Chemistry

UNIT 8 · ACIDS AND BASES

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Complete Ionization Comes First

A strong acid ionizes completely in water:

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

That reaction runs to completion, leaving no HCl molecules behind. The strong acids are HCl, HBr, HI, HNO_3, $HClO_4$ and H_2SO_4, whose first proton ionizes completely.

A strong base dissociates completely in the same way: the soluble metal hydroxides release all of their hydroxide ions, as in:

$NaOH \rightarrow Na^+ + OH^-$

Because ionization is complete, $[H_3O^+]$ or $[OH^-]$ is read straight from the recipe of the solution. No equilibrium expression and no Ka is ever written for a strong acid or base.

Weak acids and bases ionize only partially; their ion concentrations need an equilibrium calculation.

Count the Ions Each Formula Unit Releases

Monoprotic acids and monohydroxide bases count one to one: 0.010 mol L^{-1} HCl gives $[H_3O^+]$ = 0.010 mol L^{-1}, and 0.0010 mol L^{-1} NaOH gives $[OH^-]$ = 0.0010 mol L^{-1}.

Dihydroxides release two hydroxide ions per formula unit, so 0.020 mol L^{-1} $Ba(OH)_2$ has $[OH^-]$ = 0.040 mol L^{-1}, not 0.020:

$Ba(OH)_2 \rightarrow Ba^{2+} + 2OH^-$

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$Ca(OH)_2$ and $Sr(OH)_2$ behave identically.

Sulfuric acid needs the same care. Its first proton ionizes completely, and a strong base removes the second proton as well, so each H_2SO_4 molecule supplies two moles of H^+ to a neutralization:

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

Neutralization ratios are ratios of moles of H^+ to moles of OH^-, never of formula units.

The Two Definitions and the Bridge Between Them

The two definitions are:

$pH = -\log[H_3O^+]$

$pOH = -\log[OH^-]$

Both run backwards:

$[H_3O^+] = 10^{-pH}$

$[OH^-] = 10^{-pOH}$

The bridge between them comes from Kw. Taking negative logarithms of the Kw product gives:

pH + pOH = 14.00 at 25 °C

The sum is pKw, and 14.00 is its 25 °C value only.

That bridge sets the working route: acids give the pH directly and the pOH by subtraction; bases give the pOH directly and the pH = 14.00 − pOH. Every pH here is quoted to two decimal places, matching the two significant figures of most concentrations.

The Scale Is Logarithmic

One pH unit is a factor of ten in $[H_3O^+]$. A solution of pH 3 has ten times the $[H_3O^+]$ of one at pH 4 and a hundred times that of one at pH 5. A change of two pH units always means a factor of 100.

Ranking works the same way: convert to one ion before comparing. 0.10 mol L^{-1} $Ba(OH)_2$, with $[OH^-]$ = 0.20 mol L^{-1}, is more basic than 0.15 mol L^{-1} NaOH, despite its smaller formula-unit concentration.

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Dilution and Mixing Are Mole Problems

Dilution changes the volume, never the moles. Tenfold dilution of a strong acid divides $[H_3O^+]$ by ten, so the pH rises by exactly one unit: 0.10 mol L^{-1} HCl at pH 1.00 becomes 0.010 mol L^{-1} at pH 2.00, then 0.0010 mol L^{-1} at pH 3.00.

Mixing a strong acid with a strong base runs this neutralization to completion:

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

Count the moles of each ion, let the smaller amount cancel an equal amount of the other, then divide what is left by the total volume. Equal moles leave nothing over, and the mixture is neutral.

When Water Itself Enters the Calculation

Water autoionizes:

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

That reaction contributes 1.0×10^{-7} mol L^{-1} of H_3O^+. Beside 0.010 mol L^{-1} of acid that is invisible; beside 1.0×10^{-8} mol L^{-1} it is the larger share, and ignoring it gives the impossible answer pH 8 for an acid.

Charge balance keeps the books, where C is the acid concentration:

$[H_3O^+] = C + [OH^-] = C + Kw \div [H_3O^+]$

Solving gives $[H_3O^+] = 1.05 \times 10^{-7}$ mol L^{-1} for 1.0×10^{-8} mol L^{-1} HCl, so the pH is 6.98. Dilution pushes an acid's pH toward 7 from below, never past it.

Worked Examples

Example 1: A strong acid, all four quantities. For 0.050 mol L^{-1} HCl, complete ionization gives $[H_3O^+] = 0.050$ mol L^{-1}. Then pH = $-\log(0.050) = 1.30$, pOH = $14.00 - 1.30 = 12.70$, and $[OH^-] = Kw \div [H_3O^+] = (1.0 \times 10^{-14}) \div 0.050 = 2.0 \times 10^{-13}$ mol L^{-1}. Even a strongly acidic solution still contains hydroxide ions, just very few.

Example 2: A dihydroxide base. For 0.015 mol L^{-1} $Sr(OH)_2$, each formula unit releases two hydroxide ions, so $[OH^-] = 2 \times 0.015 = 0.030$ mol L^{-1}. Then pOH = $-\log(0.030) = 1.52$ and pH = $14.00 - 1.52 = 12.48$.

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Example 3: Working backwards from a pH. A solution has pH 8.70 at 25 °C. Then $[H_3O^+] = 10^{-8.70} = 2.0 \times 10^{-9}$ mol L^{-1}, pOH = 14.00 − 8.70 = 5.30, and $[OH^-] = 10^{-5.30} = 5.0 \times 10^{-6}$ mol L^{-1}. Hydroxide exceeds hydronium about 2,500-fold, which is what pH above 7 means in concentrations.

Example 4: A mixture with acid left over. Mix 50.0 mL of 0.10 mol L^{-1} HCl with 30.0 mL of 0.10 mol L^{-1} NaOH. Moles of $H^+ = 0.0500 \times 0.10 = 0.0050$ mol; moles of $OH^- = 0.0300 \times 0.10 = 0.0030$ mol. Neutralization cancels 0.0030 mol of each, leaving 0.0020 mol of H^+ in a total volume of 80.0 mL, so $[H_3O^+] = 0.0020 \div 0.0800 = 0.025$ mol L^{-1} and pH = −log(0.025) = 1.60.

Traps That Cost Marks in Topic 8.2

"0.020 M $Ba(OH)_2$, so pOH = −log(0.020) = 1.70." Each formula unit releases two hydroxide ions, so $[OH^-] = 0.040$ mol L^{-1}, pOH = 1.40 and the pH is 12.60, not 12.30.

"1.0×10^{-8} M HCl has pH 8.00, because $-\log(1.0 \times 10^{-8}) = 8$." That answer describes a basic solution, and diluting an acid can never make it basic. Water's own H_3O^+ must be counted at this concentration, and the correct pH is 6.98.

"pH + pOH = 14, always." The sum equals pKw, which is 14.00 only at 25 °C; at other temperatures the two still add to pKw, but pKw itself has moved.

"50 mL of 0.10 M H_2SO_4 is neutralized by 25 mL of 0.20 M NaOH, since 0.20 is double 0.10." Each H_2SO_4 supplies two H^+, so the acid holds 0.010 mol of H^+ and needs 0.010 mol of OH^-, which is 50 mL of 0.20 mol L^{-1} NaOH. Count moles of ions, not volumes or formula units.

"Mixing an acid of pH 1 with an equal volume of a base of pH 13 gives pH 7 because the values average out." pH values never average. Equal volumes of 0.10 mol L^{-1} HCl and 0.10 mol L^{-1} NaOH are neutral because the moles of H^+ and OH^- are exactly equal, leaving neither in excess. With any other concentrations, the leftover mole count divided by the total volume sets the pH.

"Dilute any acid tenfold and the pH rises by one unit, every time." The rule needs the acid to dominate the H_3O^+ supply. Near 1.0×10^{-7} mol L^{-1} water takes over: tenfold dilution of 1.0×10^{-7} mol L^{-1} HCl moves the pH from 6.79 only to 6.98, and no dilution carries it past 7.

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Key Relationships to Remember

- A strong acid or base ionizes completely, so $[H_3O^+]$ or $[OH^-]$ equals the stated concentration multiplied by the number of ions released per formula unit
- $pH = -\log[H_3O^+]$ and $pOH = -\log[OH^-]$
- The definitions invert as $[H_3O^+] = 10^{-pH}$ and $[OH^-] = 10^{-pOH}$
- $pH + pOH = pK_w = 14.00$ at 25 °C
- One pH unit is a factor of ten in $[H_3O^+]$; two units are a factor of 100
- $Ba(OH)_2$, $Ca(OH)_2$ and $Sr(OH)_2$ release two OH^- per formula unit; H_2SO_4 supplies two H^+ per molecule to a neutralizing base
- Dilution leaves moles unchanged: tenfold dilution moves the pH one unit toward 7 while the solute still dominates
- Mixing a strong acid and a strong base is a mole balance: subtract the smaller ion count from the larger, then divide the excess by the total volume; equal moles of H^+ and OH^- give a neutral mixture
- Below about 1×10^{-6} mol L^{-1}, water autoionization contributes significantly to $[H_3O^+]$, and an acid at 25 °C always has pH below 7, approaching 7 as the limit

Continue Revising

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Topic 8.2 Worksheet

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