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Topic 8.2: pH and pOH of Strong Acids and Bases

Worksheet · AP Chemistry

UNIT 8 · ACIDS AND BASES

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How to Use This Worksheet

Work in order. Section A checks the ideas: what complete ionization means, counting ions per formula unit, and the bridge between pH and pOH. Section B practices the working skills: pH in both directions, a dihydroxide base, dilution, mixing with an excess, and a neutralization volume. Section C is exam style: a four-solution ranking where the formula units mislead, the very dilute acid that breaks the simple rule, and a neutralization that lands exactly on balance before being pushed past it.

Reference facts. Every temperature is 25 °C, Kw = 1.0×10^{-14}, concentrations are in mol L^{-1}, and pH and pOH values are quoted to two decimal places. Strong means completely ionized throughout; no Ka and no equilibrium expression appears anywhere on this sheet, because none belongs to a strong acid or base.

One working habit ties the sheet together: name the ion you actually know before touching a logarithm. For an acid solution that ion is H_3O^+ and the pH comes first; for a base it is OH^-, the pOH comes first, and the pH follows by subtraction from 14.00. Nearly every error on this topic is a logarithm taken of the wrong ion, or of a formula-unit concentration instead of an ion concentration.

Three checks run through the whole sheet. The bridge check: the pH and pOH of one solution must add to 14.00 at 25 °C. The mole check: in any mixing problem, moles before reaction minus moles consumed equals moles left over, and the leftover concentration uses the total volume. The sanity check: at 25 °C an acid solution always has pH below 7 and a basic

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solution always has pH above 7, whatever the dilution. If a check fails, the error is in the working, never in the check.

The answer key shows every logarithm step and every mole balance in full. Do not look at the answer key until you have attempted every question in Sections A, B and C.

Section A: Recall and Concepts

A1. (a) State what the word strong means for an acid in water, in terms of what happens to its molecules. (b) Write the equation for HCl dissolving in water, and the equation for solid $Ca(OH)_2$ dissolving in water. (c) Explain why no equilibrium constant is written for either process.

A2. A solution contains 0.0010 mol L^{-1} HCl at 25 °C. Give all four quantities for this solution: $[H_3O^+]$, the pH, the pOH and $[OH^-]$.

A3. (a) State the relationship between pH and pOH that holds at any temperature, and give its numerical value at 25 °C. (b) A solution has pOH 4.60. Calculate its pH. (c) A solution has pH 3.50. Calculate $[H_3O^+]$.

A4. (a) The pH of a solution falls by 3 units. By what factor does $[H_3O^+]$ change, and in which direction? (b) The $[OH^-]$ of a solution is multiplied by 10. State what happens to its pOH and to its pH. (c) Which is more basic, a solution of pH 11.30 or a solution with $[OH^-] = 1.0 \times 10^{-2}$ mol L^{-1}? Show the conversion that decides it.

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A5. (a) A student writes "the pH of 1.0×10^{-8} mol L^{-1} HCl is 8.00". Give two distinct reasons why this answer cannot be correct. (b) State the value that the pH of an acid approaches, without reaching or passing it, as the acid is diluted again and again at 25 °C. (c) Name the second source of H_3O^+ that a correct calculation must include at such low concentrations.

Section B: Core Practice

B1. Calculate the pH, the pOH and $[OH^-]$ of a 0.025 mol L^{-1} HCl solution at 25 °C.

B2. A solution of $Ca(OH)_2$ has concentration 0.030 mol L^{-1} at 25 °C. (a) Write the dissociation equation and state $[OH^-]$. (b) Calculate the pOH. (c) Calculate the pH.

B3. A solution at 25 °C has pH 10.40. Calculate (a) $[H_3O^+]$, (b) the pOH and (c) $[OH^-]$.

B4. A 25.0 mL portion of 0.40 mol L^{-1} HNO_3 is diluted with water to a total volume of 250.0 mL. (a) Calculate the concentration of the diluted solution. (b) Calculate the pH before dilution and after dilution, and state the change in pH.

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B5. 60.0 mL of 0.050 mol L^{-1} HCl is mixed with 40.0 mL of 0.050 mol L^{-1} NaOH at 25 °C. (a) Calculate the moles of H^+ and the moles of OH^- before mixing. (b) State which ion is in excess, and by how many moles. (c) Calculate the pH of the mixture.

B6. What volume of 0.15 mol L^{-1} $Ba(OH)_2$ solution is needed to neutralize exactly 30.0 mL of 0.20 mol L^{-1} HCl? Show the mole counting in full, including the step where the two hydroxide ions per formula unit enter.

Section C: Exam Style

C1. Four solutions at 25 °C: (i) 0.10 mol L^{-1} HCl; (ii) 0.050 mol L^{-1} $Ba(OH)_2$; (iii) 0.080 mol L^{-1} NaOH; (iv) 1.0×10^{-3} mol L^{-1} HNO_3. (a) Calculate the pH of each solution. (b) Rank the four from lowest pH to highest pH. (c) A student ranks (ii) below (iii) in basicity because 0.050 is less than 0.080. Explain, using the ion counts, why the student's order is wrong.

C2. A solution of HCl has concentration 1.0×10^{-7} mol L^{-1} at 25 °C. (a) Write the charge balance connecting $[H_3O^+]$, the acid concentration and $[OH^-]$, and combine it with Kw to obtain a quadratic equation in $[H_3O^+]$. (b) Solve the quadratic and calculate the pH. (c) State the pH given by the simple rule $[H_3O^+]$ = acid concentration, and explain in one sentence why that rule fails here. (d) A 1.0×10^{-7} mol L^{-1} NaOH solution has pH 7.21. Explain why this value and your part (b) answer sit symmetrically either side of 7.00.

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C3. 100 mL of 0.20 mol L^{-1} HCl is mixed with 100 mL of 0.10 mol L^{-1} $Ba(OH)_2$ at 25 °C. (a) Calculate the moles of H^+ and the moles of OH^-, and hence the pH of the mixture. (b) The experiment is repeated using 100 mL of 0.30 mol L^{-1} HCl against the same $Ba(OH)_2$ solution. Calculate the pH now. (c) State the volume used as the divisor in part (b), and why it, rather than either starting volume, is correct.

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Answer Key

A1. (a) Strong means every acid molecule ionizes: the fraction ionized is effectively 100 percent, so no intact HCl molecules remain in the solution. (b) $HCl + H_2O \rightarrow H_3O^+ + Cl^-$, and $Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$. Both arrows point one way. (c) An equilibrium constant describes a balance between forward and reverse processes. Here there is no measurable reverse process and no equilibrium mixture of molecules and ions to describe: the reactant is consumed completely, so the ion concentrations follow from the starting concentration by counting, not from an equilibrium expression. Counting replaces equilibrium here: the starting concentration and the ion ratio in the formula fix everything that follows.

A2. Complete ionization gives $[H_3O^+] = 0.0010$ mol L^{-1} directly. $pH = -\log(1.0 \times 10^{-3}) = 3.00$. The bridge gives $pOH = 14.00 - 3.00 = 11.00$. Then $[OH^-] = Kw \div [H_3O^+] = (1.0 \times 10^{-14}) \div (1.0 \times 10^{-3}) = 1.0 \times 10^{-11}$ mol L^{-1}. The four answers are consistent: 3.00 + 11.00 = 14.00, and $(1.0 \times 10^{-3})(1.0 \times 10^{-11}) = 1.0 \times 10^{-14}$.

A3. (a) pH + pOH = pKw at any temperature; at 25 °C, pKw = 14.00. (b) $pH = 14.00 - 4.60 = 9.40$. (c) $[H_3O^+] = 10^{-3.50} = 3.2 \times 10^{-4}$ mol L^{-1}. A pH halfway between two whole units gives a concentration between the two powers of ten, which is the signature of the 10^{-pH} inversion: the .50 in the exponent produces the factor 3.2, not 1.0.

A4. (a) $[H_3O^+]$ increases by a factor of $10^3 = 1{,}000$. Each unit of pH fall is a tenfold rise, and three units multiply as 10 × 10 × 10. (b) The pOH falls by exactly 1, because pOH is the negative log of the quantity that grew tenfold. The pH therefore rises by exactly 1, since the two must still add to 14.00 at 25 °C. (c) Convert the hydroxide concentration: $pOH = -\log(1.0 \times 10^{-2}) = 2.00$, so its $pH = 14.00 - 2.00 = 12.00$. That exceeds 11.30, so the solution with $[OH^-] = 1.0 \times 10^{-2}$ mol L^{-1} is the more basic. The two givens could not be compared until both were expressed as pH.

A5. (a) First, 8.00 is above 7, which describes a basic solution; an acid, however dilute, cannot cross to the basic side at 25 °C. Second, the calculation counts only the H_3O^+ from the acid and omits the H_3O^+ from the autoionization of water, which at a concentration of 1.0×10^{-8} mol L^{-1} is actually the bigger source: water alone holds 1.0×10^{-7} mol L^{-1}. (b) 7. (c) The autoionization of water, $2H_2O \rightleftharpoons H_3O^+ + OH^-$.

B1. Complete ionization: $[H_3O^+] = 0.025$ mol L^{-1}. $pH = -\log(0.025) = 1.60$. $pOH = 14.00 - 1.60 = 12.40$. $[OH^-] = (1.0 \times 10^{-14}) \div 0.025 = 4.0 \times 10^{-13}$ mol L^{-1}. Bridge check: 1.60 + 12.40 = 14.00.

B2. (a) $Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$. Each formula unit releases two hydroxide ions, so $[OH^-] = 2 \times 0.030 = 0.060$ mol L^{-1}. (b) $pOH = -\log(0.060) = 1.22$. (c) $pH = 14.00 - 1.22 = 12.78$. The single most common error in this topic is skipping the factor of two in part (a) and landing on 12.48; the formula unit count is where the marks are decided.

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B3. (a) $[H_3O^+] = 10^{-10.40} = 4.0 \times 10^{-11}$ mol L^{-1}. (b) pOH = 14.00 − 10.40 = 3.60. (c) $[OH^-] = 10^{-3.60} = 2.5 \times 10^{-4}$ mol L^{-1}. Note the cross-check available in both directions: part (c) can equally be reached from Kw, $(1.0 \times 10^{-14}) \div (4.0 \times 10^{-11}) = 2.5 \times 10^{-4}$ mol L^{-1}. The solution is basic, and its hydroxide concentration is the larger of the two, as a pH above 7 requires.

B4. (a) Dilution keeps the moles fixed: moles = 0.0250 × 0.40 = 0.010 mol, now in 0.2500 L, so the concentration is 0.010 ÷ 0.2500 = 0.040 mol L^{-1}. Equivalently, the volume grew tenfold, so the concentration fell tenfold. (b) Before: pH = −log(0.40) = 0.40. After: pH = −log(0.040) = 1.40. The change is +1.00, exactly the one unit a tenfold dilution always produces while the acid still dominates the H_3O^+ supply. A hundredfold dilution would add a second unit, giving pH 2.40: each power of ten in the volume is worth exactly one unit on the scale.

B5. (a) Moles of H^+ = 0.0600 × 0.050 = 0.0030 mol. Moles of OH^- = 0.0400 × 0.050 = 0.0020 mol. (b) H^+ is in excess by 0.0030 − 0.0020 = 0.0010 mol; the neutralization $H_3O^+ + OH^- \rightarrow 2H_2O$ consumes all 0.0020 mol of OH^- and an equal amount of H^+. (c) The leftover 0.0010 mol of H^+ sits in the total volume, 60.0 + 40.0 = 100.0 mL = 0.1000 L, so $[H_3O^+] = 0.0010 \div 0.1000 = 0.010$ mol L^{-1} and pH = −log(0.010) = 2.00.

B6. Count the H^+ first: moles of H^+ = 0.0300 × 0.20 = 0.0060 mol, and neutralization needs the same number of moles of OH^-, 0.0060 mol. Each $Ba(OH)_2$ formula unit supplies two OH^-, so the moles of $Ba(OH)_2$ required are 0.0060 ÷ 2 = 0.0030 mol. Volume = moles ÷ concentration = 0.0030 ÷ 0.15 = 0.0200 L = 20.0 mL. The division by two is the step that carries the dihydroxide stoichiometry; missing it doubles the answer to 40.0 mL.

C1. (a) (i) $[H_3O^+] = 0.10$ mol L^{-1}, pH = −log(0.10) = 1.00. (ii) $[OH^-] = 2 \times 0.050 = 0.10$ mol L^{-1}, pOH = 1.00, pH = 14.00 − 1.00 = 13.00. (iii) $[OH^-] = 0.080$ mol L^{-1}, pOH = −log(0.080) = 1.10, pH = 14.00 − 1.10 = 12.90. (iv) $[H_3O^+] = 1.0 \times 10^{-3}$ mol L^{-1}, pH = 3.00. (b) Lowest to highest: (i) 1.00, (iv) 3.00, (iii) 12.90, (ii) 13.00. Note how close (ii) and (iii) sit, just 0.10 apart, even though their hydroxide concentrations differ by a factor of 1.25: on a logarithmic scale a factor of 1.25 is a small step, which is why the ion count, not the pH gap, is the reliable thing to compare first. (c) Solution (ii) holds fewer formula units per liter than (iii), 0.050 against 0.080, but each $Ba(OH)_2$ unit releases two hydroxide ions against NaOH's one, so (ii) has the larger hydroxide concentration, 0.10 against 0.080 mol L^{-1}. Basicity is set by $[OH^-]$, not by the concentration of formula units, so (ii) is the more basic and stands above (iii) in the ranking.

C2. (a) The solution must be electrically neutral, so the positive charges balance: $[H_3O^+] = [Cl^-] + [OH^-] = C + [OH^-]$, with $C = 1.0 \times 10^{-7}$ mol L^{-1}. Substituting $[OH^-] = Kw \div [H_3O^+]$ and clearing the fraction gives $[H_3O^+]^2 - C[H_3O^+] - Kw = 0$, that is $x^2 - (1.0 \times 10^{-7})x - 1.0 \times 10^{-14} = 0$. (b) $x = [1.0 \times 10^{-7} + \sqrt{((1.0 \times 10^{-7})^2 + 4 \times 1.0 \times 10^{-14})}] \div 2 = [1.0 \times 10^{-7} + 2.24 \times 10^{-7}] \div 2 = 1.62 \times 10^{-7}$ mol L^{-1}, taking the positive root. pH = −log(1.62×10^{-7}) = 6.79. (c) The simple rule gives $[H_3O^+] = 1.0 \times 10^{-7}$ mol L^{-1} and pH 7.00, the value for pure water; it fails because at this concentration the H_3O^+ supplied by water itself is the same size as the H_3O^+ supplied by the acid, so neither source can be dropped. (d)

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The base case is the mirror image: $[OH^-]$ obeys the identical quadratic with the identical numbers, giving $[OH^-] = 1.62 \times 10^{-7}$ mol L^{-1}, pOH 6.79 and pH 7.21. Kw treats the two ions symmetrically, so the excess of one ion over 1.0×10^{-7} in the acid case equals the excess in the base case, and the two pH values sit 0.21 on either side of 7.00.

C3. (a) Moles of $H^+ = 0.100 \times 0.20 = 0.020$ mol. The $Ba(OH)_2$ portion holds $0.100 \times 0.10 = 0.010$ mol of formula units, each releasing two OH^-, so moles of $OH^- = 0.020$ mol. The two counts are exactly equal, nothing is left over, and the mixture is neutral: pH = 7.00. (b) Moles of $H^+ = 0.100 \times 0.30 = 0.030$ mol, against the same 0.020 mol of OH^-, leaving 0.010 mol of H^+ in excess. That excess sits in 200 mL = 0.200 L of mixture, so $[H_3O^+] = 0.010 \div 0.200 = 0.050$ mol L^{-1} and $pH = -\log(0.050) = 1.30$. (c) The divisor is the total volume, 0.200 L. Concentration is defined per liter of the final mixture; the leftover ions are spread through all of it, not confined to the portion they arrived in. Using 0.100 L would double the concentration and shave 0.30 off the pH.

Continue Revising

Topic 8.2 Revision Notes

Topic 8.2 Quiz: pH and pOH of Strong Acids and Bases

Previous Worksheet: Topic 8.1 Introduction to Acids and Bases

Next Worksheet: Topic 8.3 Weak Acid and Base Equilibria

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