Topic 8.3: Weak Acid and Base Equilibria

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: writing Ka and Kb expressions, ranking by constant size, conjugate conversions through Kw, and reading salt solutions. Section B practices the working skills: ICE tables for a weak acid and a weak base, a constant recovered from a pH in each direction, a percent ionization that breaks the approximation, and a dilution pair. Section C is exam style: a basic salt solution built from a conjugate Kb, two acids whose constants differ a hundredfold, and an acidic salt whose constant must be earned through Kw first.

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. Each question supplies the constant it needs. Two conjugate pairs recur throughout: Ka of CH_3COOH = 1.8×10^{-5} with Kb of CH_3COO^- = 5.6×10^{-10}, and Kb of NH_3 = 1.8×10^{-5} with Ka of NH_4^+ = 5.6×10^{-10}.

One working habit ties the sheet together: never let an approximation leave the page unchecked. Every ICE calculation ends by computing percent ionization, (x ÷ C) × 100 percent. At 5 percent or below the small-x route stands; above it the quadratic replaces it, and the answer key states which route each question needed and why.

Two checks run through the whole sheet. The pair check: the Ka and Kb of a conjugate pair multiply to 1.0×10^{-14}, and their pKa and pKb add to 14.00. The round-trip check: the x you report, substituted back into $x^2 \div (C - x)$, returns the constant you were given. If either check fails, the error is in the working, never in the check.

The sheet also assumes that the ionization being studied is the only equilibrium that matters. Water's own autoionization is ignored beside the x supplied by the acid or base, which is safe

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whenever x is far above 10^{-7} mol L^{-1}, as it is in every question here. Square brackets always mean equilibrium concentrations.

The answer key shows every ICE row, every logarithm step and every check 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) Write the equation for HF ionizing in water, and the expression for its Ka. (b) Write the equation for NH_3 reacting with water, and the expression for its Kb. (c) State what a small Ka value tells you about where the ionization equilibrium lies and about the fraction of molecules ionized.

A2. Five acids and their Ka values: HF 6.8×10^{-4}, HCOOH 1.8×10^{-4}, CH_3COOH 1.8×10^{-5}, H_2CO_3 4.3×10^{-7}, HCN 4.9×10^{-10}. (a) Rank the acids from strongest to weakest. (b) State which acid has the largest pKa, and calculate that pKa. (c) State the general link between the size of pKa and acid strength.

A3. (a) The Ka of an acid HA is 1.0×10^{-5}. Calculate the Kb of its conjugate base A^-. (b) The Kb of NH_3 is 1.8×10^{-5}. Calculate the Ka of NH_4^+. (c) For CH_3COOH ($Ka = 1.8 \times 10^{-5}$), calculate the pKa, and then the pKb of CH_3COO^- using the pair relation.

A4. (a) Define percent ionization for a weak acid of concentration C that produces $[H_3O^+] = x$ at equilibrium. (b) A 0.10 mol L^{-1} weak acid has $[H_3O^+] = 1.3 \times 10^{-3}$ mol L^{-1} at equilibrium. Calculate its percent ionization. (c) The same acid is diluted to 0.0010 mol L^{-1}. State what happens to the percent ionization and what happens to $[H_3O^+]$, and name the principle that explains the change in percent ionization.

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A5. Predict whether each solution is acidic, basic or neutral, justifying each answer by naming the ion that reacts with water, if any: (a) $NaCH_3COO$, (b) NH_4Cl, (c) KNO_3, (d) Na_2CO_3.

Section B: Core Practice

B1. Formic acid, HCOOH, has Ka = 1.8×10^{-4}. Calculate the pH of a 0.20 mol L^{-1} solution, showing the ICE rows, the small-x approximation, its percent check, and the quadratic value of x.

B2. Calculate the pH of a 0.050 mol L^{-1} solution of NH_3 (Kb = 1.8×10^{-5}), showing the pOH step explicitly.

B3. A 0.10 mol L^{-1} solution of a weak acid has pH 3.30. Calculate the Ka of the acid and its percent ionization.

B4. A weak acid HA has Ka = 2.0×10^{-4}. Calculate the percent ionization of HA in a 0.050 mol L^{-1} solution. Show the approximation result as well as the value you actually report, and state why they differ.

B5. A 0.10 mol L^{-1} solution of a weak base B has pH 9.80. Calculate the Kb of B.

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B6. CH_3COOH has Ka = 1.8×10^{-5}. (a) Calculate the pH and the percent ionization of a 0.10 mol L^{-1} solution, using the quadratic. (b) Repeat the calculation for a 0.0010 mol L^{-1} solution. (c) Using your numbers, state why the small-x approximation is acceptable in one case and not in the other.

Section C: Exam Style

C1. Sodium acetate, $NaCH_3COO$, dissolves completely to give Na^+ and CH_3COO^- ions. Ka of CH_3COOH is 1.8×10^{-5}. (a) Write the equation for the ion that reacts with water, and explain why Na^+ plays no part. (b) Calculate the Kb of that ion. (c) Calculate the pH of a 0.010 mol L^{-1} solution of $NaCH_3COO$. (d) Calculate the fraction of the ion that has reacted at equilibrium, and comment on the approximation used in part (c).

C2. Two weak acids, P (Ka = 1.0×10^{-4}) and Q (Ka = 1.0×10^{-6}), are both made up at 0.10 mol L^{-1}. (a) Calculate $[H_3O^+]$ and the pH of each solution. (b) Calculate the ratio of the two $[H_3O^+]$ values. (c) Explain why a hundredfold difference in Ka does not produce a hundredfold difference in $[H_3O^+]$.

C3. Ammonium chloride, NH_4Cl, dissolves completely to give NH_4^+ and Cl^- ions. Kb of NH_3 is 1.8×10^{-5}. (a) Calculate the Ka of NH_4^+. (b) Calculate the pH of a 0.10 mol L^{-1} solution of NH_4Cl. (c) Explain why this solution is acidic even though the salt contains the anion of a strong acid.

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Answer Key

A1. (a) $HF + H_2O \rightleftharpoons H_3O^+ + F^-$, with $K_a = [H_3O^+][F^-] \div [HF]$. (b) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$, with $K_b = [NH_4^+][OH^-] \div [NH_3]$. Water appears in neither expression: it is the solvent, and its concentration is folded into the constant. Every concentration in both expressions is raised to its stoichiometric coefficient, which is 1 throughout this topic. (c) The equilibrium lies far to the left: only a small fraction of the molecules are ionized at any moment, and most of the acid remains as intact HA.

A2. (a) HF > HCOOH > CH_3COOH > H_2CO_3 > HCN. At equal concentration the larger Ka sits further right and gives the stronger acid. (b) HCN, the weakest acid: $pK_a = -\log(4.9 \times 10^{-10}) = 9.31$. (c) pKa runs backwards against strength: the larger the pKa, the smaller the Ka and the weaker the acid.

A3. (a) $K_b = K_w \div K_a = (1.0 \times 10^{-14}) \div (1.0 \times 10^{-5}) = 1.0 \times 10^{-9}$. (b) $K_a = K_w \div K_b = (1.0 \times 10^{-14}) \div (1.8 \times 10^{-5}) = 5.6 \times 10^{-10}$. (c) $pK_a = -\log(1.8 \times 10^{-5}) = 4.74$. Then $pK_b = 14.00 - 4.74 = 9.26$ by the pair relation, matching Kb of CH_3COO^- = 5.6×10^{-10}. The sum 4.74 + 9.26 = 14.00 is the pair check on the two values.

A4. (a) Percent ionization = (x ÷ C) × 100 percent: the concentration that has ionized, expressed as a percentage of the starting concentration. (b) $(1.3 \times 10^{-3} \div 0.10) \times 100 = 1.3$ percent. (c) The percent ionization increases, while $[H_3O^+]$ decreases, so the pH rises toward 7. By Le Chatelier's principle, dilution shifts the equilibrium toward the side with more dissolved particles, the ions, so a greater fraction of the remaining molecules ionizes even though there are fewer of them per liter. The two changes pull in opposite directions, which is exactly why percent ionization and acidity must never be treated as the same quantity.

A5. (a) Basic. CH_3COO^- is the conjugate base of the weak acid CH_3COOH and reacts with water to produce OH^-; Na^+ is a spectator. (b) Acidic. NH_4^+ is the conjugate acid of the weak base NH_3 and donates a proton to water; Cl^- is a spectator. (c) Neutral. K^+ comes from the strong base KOH and NO_3^- from the strong acid HNO_3, so neither ion reacts with water. (d) Basic. CO_3^{2-} is the conjugate base of the weak acid HCO_3^- and reacts: $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$; Na^+ is a spectator.

B1. ICE rows for $HCOOH + H_2O \rightleftharpoons H_3O^+ + HCOO^-$: initial 0.20, 0, 0; change −x, +x, +x; equilibrium 0.20 − x, x, x, with $x = [H_3O^+]$. The equilibrium condition is $K_a = x^2 \div (0.20 - x) = 1.8 \times 10^{-4}$. Approximation: $x \approx \sqrt{(1.8 \times 10^{-4} \times 0.20)} = \sqrt{(3.6 \times 10^{-5})} = 6.0 \times 10^{-3}$ mol L^{-1}. Percent ionization = $6.0 \times 10^{-3} \div 0.20 \times 100 = 3.0$ percent, inside the check, with an error of about 1.5 percent in x. Quadratic: $x = 5.9 \times 10^{-3}$ mol L^{-1}. Round trip: $(5.9 \times 10^{-3})^2 \div (0.20 - 0.0059)$ returns 1.8×10^{-4}. $pH = -\log(5.9 \times 10^{-3}) = 2.23$.

B2. ICE rows for $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ give $x = [OH^-] \approx \sqrt{(1.8 \times 10^{-5} \times 0.050)} = \sqrt{(9.0 \times 10^{-7})} = 9.5 \times 10^{-4}$ mol L^{-1}. Percent ionization = 1.9 percent, so the check passes; the quadratic gives $x = 9.4 \times 10^{-4}$ mol L^{-1}. Then $pOH = -\log(9.4 \times 10^{-4}) = 3.03$ and $pH = 14.00 - 3.03 = 10.97$. Round trip: $(9.4 \times 10^{-4})^2$

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÷ (0.050 − 0.00094) returns 1.8×10^{-5}. The pH is reached only through the pOH: taking −log of $[OH^-]$ and reporting it as a pH is the standing error in base problems.

B3. The pH fixes x directly: $x = [H_3O^+] = 10^{-3.30} = 5.0 \times 10^{-4}$ mol L^{-1}. Then $K_a = x^2 \div (C - x) = (5.0 \times 10^{-4})^2 \div (0.10 - 0.0005) = 2.5 \times 10^{-6}$. Percent ionization = $5.0 \times 10^{-4} \div 0.10 \times 100 = 0.50$ percent. No quadratic is needed in this direction: a measured pH is already an equilibrium value.

B4. Approximation: $x \approx \sqrt{(2.0 \times 10^{-4} \times 0.050)} = \sqrt{(1.0 \times 10^{-5})} = 3.2 \times 10^{-3}$ mol L^{-1}, which would give a percent ionization of 6.3 percent. That is above the 5 percent line, so the approximation is not accepted. Quadratic: $x = 3.1 \times 10^{-3}$ mol L^{-1}, and the percent ionization to report is $3.1 \times 10^{-3} \div 0.050 \times 100 = 6.1$ percent. The approximation overstates x by about 3 percent because the amount ionized is a real share of 0.050 mol L^{-1}, so replacing C − x by C is no longer harmless.

B5. First the pOH: pOH = 14.00 − 9.80 = 4.20. Then $x = [OH^-] = 10^{-4.20} = 6.3 \times 10^{-5}$ mol L^{-1}. So $K_b = x^2 \div (C - x) = (6.3 \times 10^{-5})^2 \div (0.10 - 0.000063) = 4.0 \times 10^{-8}$. The percent ionization is $6.3 \times 10^{-5} \div 0.10 \times 100 = 0.063$ percent. The subtraction in the denominator changes nothing at this precision, but the route matters: pH to pOH to $[OH^-]$ to Kb, never pH straight into the Kb expression.

B6. (a) Quadratic: $x^2 + (1.8 \times 10^{-5})x - 1.8 \times 10^{-6} = 0$ gives $x = 1.33 \times 10^{-3}$ mol L^{-1}, so pH = 2.88 and percent ionization = 1.33 percent. Round trip: $(1.33 \times 10^{-3})^2 \div (0.10 - 0.00133)$ returns 1.8×10^{-5}. The approximation gives $x = 1.34 \times 10^{-3}$ mol L^{-1}, an error of 0.7 percent. (b) Quadratic: $x = 1.25 \times 10^{-4}$ mol L^{-1}, so pH = 3.90 and percent ionization = 12.5 percent. (c) At 0.10 mol L^{-1} only 1.33 percent ionizes, below the 5 percent line, so dropping x beside C costs under 1 percent in x and the approximation is acceptable. At 0.0010 mol L^{-1} the ionized share is 12.5 percent, x is a real part of C, and the approximation returns 1.34×10^{-4} mol L^{-1}, about 7 percent too large: only the quadratic is honest there. Dilution raised the percent ionized roughly tenfold while $[H_3O^+]$ fell roughly tenfold.

C1. (a) $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$. Na^+ is the cation of the strong base NaOH; it has no proton to give and no tendency to take one, so it is a spectator and the acetate ion alone sets the pH. (b) $K_b = K_w \div K_a = (1.0 \times 10^{-14}) \div (1.8 \times 10^{-5}) = 5.6 \times 10^{-10}$. (c) $x = [OH^-] \approx \sqrt{(5.6 \times 10^{-10} \times 0.010)} = \sqrt{(5.6 \times 10^{-12})} = 2.4 \times 10^{-6}$ mol L^{-1}. pOH = −log(2.4×10^{-6}) = 5.63, and pH = 14.00 − 5.63 = 8.37. (d) Fraction reacted = $2.4 \times 10^{-6} \div 0.010 = 2.4 \times 10^{-4}$, about 0.024 percent. So little of the ion reacts that neglecting x beside 0.010 in the denominator changes nothing at the precision quoted: the approximation in part (c) is fully justified.

C2. (a) Acid P: approximation $x \approx \sqrt{(1.0 \times 10^{-4} \times 0.10)} = 3.2 \times 10^{-3}$ mol L^{-1}, a percent ionization of 3.2 percent, which passes the check; the quadratic gives $x = 3.1 \times 10^{-3}$ mol L^{-1} and pH = 2.51. Acid Q: $x \approx \sqrt{(1.0 \times 10^{-6} \times 0.10)} = 3.2 \times 10^{-4}$ mol L^{-1}, 0.32 percent ionized, and the quadratic agrees at this precision; pH = 3.50. (b) Ratio = $3.1 \times 10^{-3} \div 3.2 \times 10^{-4} \approx 10$ (9.9 from the unrounded roots). (c) At the same concentration, $x \approx \sqrt{(K_a \times C)}$, so $[H_3O^+]$ scales with the square root of Ka. A factor of 100 in Ka becomes $\sqrt{100} = 10$ in $[H_3O^+]$, which is why the pH values sit about one unit apart, 2.51 against 3.50, rather than two. The square root is already visible in $K_a = x^2 \div (C - x)$: doubling $[H_3O^+]$

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at fixed C would quadruple the numerator, so the constant, not the concentration, carries the full hundredfold factor.

C3. (a) Ka of NH_4^+ = Kw ÷ Kb of NH_3 = $(1.0 \times 10^{-14}) \div (1.8 \times 10^{-5}) = 5.6 \times 10^{-10}$. (b) For $NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH_3$, $x = [H_3O^+] \approx \sqrt{(5.6 \times 10^{-10} \times 0.10)} = \sqrt{(5.6 \times 10^{-11})} = 7.5 \times 10^{-6}$ mol L^{-1}. The percent ionization is 0.0075 percent, so the approximation is safe. $pH = -\log(7.5 \times 10^{-6}) = 5.13$. Round trip: $(7.5 \times 10^{-6})^2 \div (0.10 - 0.0000075)$ returns 5.6×10^{-10}, the Ka earned in part (a). (c) The acidity comes from the cation, not the anion. NH_4^+ is the conjugate acid of the weak base NH_3, so it donates protons to water and produces H_3O^+; Cl^-, the conjugate base of strong HCl, does not react with water at all. A salt solution is judged ion by ion, never by the salt's name.

Continue Revising

Topic 8.3 Revision Notes

Topic 8.3 Quiz: Weak Acid and Base Equilibria

Previous Worksheet: Topic 8.2 pH and pOH of Strong Acids and Bases

Next Worksheet: Topic 8.4 Acid-Base Reactions and Buffers

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