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Topic 4.8: Introduction to Acid-Base Reactions

Worksheet · AP Chemistry

UNIT 4 · CHEMICAL REACTIONS

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How to Use This Worksheet

Work in order. Section A checks the definitions, conjugate pairs, and the strong/weak distinction. Section B practises net ionic equations for neutralisations, carbonate accepting protons one step at a time, and reading a particle snapshot. Section C combines stepwise proton donation, a measured strength comparison, and amphiprotic behaviour in exam style.

Reference facts used throughout. The strong acids are HCl, HBr, HI, HNO_3, $HClO_4$, and the first proton of H_2SO_4. The strong bases are the Group 1 hydroxides and $Ba(OH)_2$. A single arrow, →, means a transfer or ionisation goes to completion; ⇌ means an equilibrium in which most particles stay unreacted. $H^+(aq)$ is shorthand for $H_3O^+(aq)$. The molar volume of a gas at STP is 22.4 L/mol. Carry unrounded values through each calculation and round only the final answer, to three significant figures.

Three tools recur in nearly every question. Build a conjugate by adding or removing exactly one H^+. Build a net ionic equation by writing what is actually present, strong acids and bases as ions and weak species whole, then deleting the spectators. Judge strength by the fraction of particles that have transferred a proton, never by the amount dissolved.

Do not look at the answer key until you have attempted every question.

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Section A: Recall and Concepts

A1. Define each term in one sentence. (a) Brønsted-Lowry acid. (b) Brønsted-Lowry base. (c) Conjugate acid-base pair. (d) Amphiprotic species. (e) Arrhenius base, and why NH_3 counts as a base under Brønsted-Lowry but not under Arrhenius. (f) Neutralisation, defined at the particle level.

A2. Write the conjugate base of each species. (a) H_2O (b) HCO_3^- (c) NH_4^+ (d) HSO_4^- (e) H_3PO_4

A3. Write the conjugate acid of each species. (a) OH^- (b) CO_3^{2-} (c) CH_3COO^- (d) H_2O (e) SO_4^{2-}

A4. For each equation, identify the acid, the base, the conjugate acid and the conjugate base. (a) $HF + H_2O \rightleftharpoons H_3O^+ + F^-$ (b) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ (c) $HCO_3^- + OH^- \rightarrow CO_3^{2-} + H_2O$ (d) $HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$ (e) $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

A5. Each solution is 0.10 mol/L. For each solute, state whether it is a strong or weak acid or base, and name the main particles present, other than water molecules. (a) HCl (b) CH_3COOH (c) NaOH (d) NH_3 (e) HNO_3 (f) $Ba(OH)_2$

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Section B: Core Practice

B1. Hydrochloric acid is added to potassium hydroxide solution. (a) Write the balanced molecular equation. (b) Write the net ionic equation. (c) Name the spectator ions. (d) Name the salt formed. (e) A 25.0 mL portion of the potassium hydroxide solution contains 0.00500 mol of OH^- ions. Calculate the moles of HCl needed to neutralise this portion exactly.

B2. Sodium hydroxide solution is added to acetic acid solution. (a) Write the balanced molecular equation. (b) Write the net ionic equation. (c) Explain why CH_3COOH is written whole in the net ionic equation while OH^- is written as an ion. (d) Name the spectator ion.

B3. Ammonia solution is mixed with hydrochloric acid. (a) Write the balanced molecular equation. (b) Write the net ionic equation. (c) Explain why neither OH^- nor H_2O appears in the net ionic equation, and why the reaction still counts as a neutralisation. (d) Name the spectator ion.

B4. A solution contains 0.0200 mol of CO_3^{2-} ions, and hydrochloric acid is added to it. (a) Acid is added until exactly 0.0200 mol of H^+ has been supplied. Write the net ionic equation for the change, and state what the solution now contains and how much of it. (b) A further 0.0200 mol of H^+ is added. Write the net ionic equation for this second change. (c) Calculate the volume of gas produced in part (b), measured at STP. (d) Write the overall net ionic equation from CO_3^{2-} to the final products.

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B5. A flask contains 0.0150 mol of dissolved HCl. (a) State how many moles of H_3O^+ are present, and justify your answer. (b) Calculate the moles of NaOH needed to neutralise the acid exactly. (c) Calculate the volume of 0.150 mol/L NaOH solution that contains this amount. (d) The flask instead contains 0.0150 mol of CH_3COOH. State the moles of NaOH needed now, and explain why the answer is unchanged even though acetic acid is weak. (e) Before neutralisation, the 0.0150 mol of HCl from part (a) is made up to a total volume of 250 mL with water. Calculate the acid concentration, and state whether the moles of NaOH needed in part (b) change.

B6. A snapshot of one small region of an ammonia solution shows 8 NH_3 molecules and 8 H_2O molecules. A later snapshot of the same region shows 6 NH_3, 6 H_2O, 2 NH_4^+ and 2 OH^-. (a) Write the equation for the change. (b) Identify the acid and the base. (c) Identify the two conjugate pairs. (d) Explain what the snapshot shows about ammonia as a base. (e) Describe the later snapshot if the base were strong instead, with the same starting counts. (f) Calculate the fraction of ammonia molecules that have accepted a proton in the later snapshot, as a percentage.

Section C: Exam Style

C1. Sulfuric acid is diprotic: it donates its two protons in separate steps. (a) Write the equation for the transfer of its first proton to water. (b) State the conjugate base of H_2SO_4. (c) Write the equation for the transfer of the second proton to water. Use the correct arrow, and justify your choice. (d) HSO_4^- can also accept a proton. Write the equation for HSO_4^- reacting with H_3O^+, and state the term for a species that behaves as HSO_4^- does across parts (c) and (d). (e) A student writes that the conjugate base of H_2SO_4 is SO_4^{2-}. Explain the error.

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C2. Two acid solutions, both 0.10 mol/L, are compared: hydrochloric acid and acetic acid. Measurement shows that the hydronium ion concentration in the acetic acid is 1.3×10^{-3} mol/L. (a) State the hydronium ion concentration in the hydrochloric acid, and explain why it takes that value. (b) Calculate how many times larger the hydronium ion concentration in the hydrochloric acid is. (c) Calculate the percentage of acetic acid molecules that are ionised. (d) Equal volumes of the two solutions are separately neutralised with NaOH solution. State, with a reason, which solution needs more NaOH. (e) Write the net ionic equation for each neutralisation. (f) A student argues that the acetic acid must need less NaOH, because its hydronium concentration is 77 times smaller. State the flaw in the argument in one sentence.

C3. Consider the equilibrium $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$. (a) Identify the two conjugate pairs. (b) State which species is the strongest acid present, and use your answer to explain why the equilibrium lies to the left. (c) In $HCl + H_2O \rightarrow H_3O^+ + Cl^-$, water accepts a proton; in the equilibrium above, water donates one. State the term for this behaviour of water, and explain in one sentence how the same substance can play both roles. (d) A solution of NH_4Cl is found to contain more H_3O^+ than pure water does. Explain this observation in terms of proton transfer. (e) Ammonia solution is now added to an acidic solution, so NH_3 meets H_3O^+ directly. Write the equation for this proton transfer, and identify the acid and the base.

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Answer Key

A1. (a) A species that donates a proton (H^+) to another species. (b) A species that accepts a proton (H^+) from another species. (c) Two species that differ by exactly one proton, one being the acid form and the other the base form of the same species. (d) A species that can both donate and accept a proton, acting as an acid in one reaction and as a base in another. (e) A substance that produces OH^- ions in water. NH_3 has no hydroxide to release, so it fails the Arrhenius test; it is a Brønsted-Lowry base because it accepts a proton from water, producing OH^- by the transfer itself. (f) The transfer of a proton from an acid to a base.

A2. Remove one H^+ each time; the charge falls by one. (a) OH^-. (b) CO_3^{2-}. (c) NH_3. (d) SO_4^{2-}. (e) $H_2PO_4^-$.

A3. Add one H^+ each time; the charge rises by one. (a) H_2O. (b) HCO_3^-. (c) CH_3COOH. (d) H_3O^+. (e) HSO_4^-.

A4. (a) Acid: HF; base: H_2O; conjugate acid: H_3O^+; conjugate base: F^-. (b) Acid: H_2O; base: NH_3; conjugate acid: NH_4^+; conjugate base: OH^-. (c) Acid: HCO_3^-; base: OH^-; conjugate acid: H_2O; conjugate base: CO_3^{2-}. (d) Acid: HSO_4^-; base: H_2O; conjugate acid: H_3O^+; conjugate base: SO_4^{2-}. (e) Acid: CH_3COOH; base: OH^-; conjugate acid: H_2O; conjugate base: CH_3COO^-.

A5. (a) Strong acid: the particles are H_3O^+ and Cl^- ions; no HCl molecules remain. (b) Weak acid: mostly CH_3COOH molecules, with small amounts of H_3O^+ and CH_3COO^- ions. (c) Strong base: Na^+ and OH^- ions only. (d) Weak base: mostly NH_3 molecules, with small amounts of NH_4^+ and OH^- ions. (e) Strong acid: H_3O^+ and NO_3^- ions; no HNO_3 molecules remain. (f) Strong base: Ba^{2+} and OH^- ions only, with twice as many OH^- ions as Ba^{2+} ions.

B1. (a) $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$. (b) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. (c) K^+ and Cl^-. (d) Potassium chloride. (e) The net ionic equation runs 1 : 1, so moles of HCl = moles of OH^- = 0.00500 mol.

B2. (a) $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$. (b) $CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)$. (c) Only a very small fraction of acetic acid molecules are ionised, so the species actually present in the flask is the intact CH_3COOH molecule, and it must be written whole. The OH^- comes from sodium hydroxide, a strong base that is fully ionised, so it is present as free ions and written as OH^-. (d) Na^+.

B3. (a) $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$. (b) $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$. (c) The ammonia takes the proton directly from the acid, so no hydroxide ions take part and no water is formed: the proton simply changes hands. It is still a neutralisation because a proton is transferred from an acid to a base, which is the defining event; forming water is only what happens when the base happens to be OH^-. (d) Cl^-.

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B4. (a) $CO_3^{2-}(aq) + H^+(aq) \rightarrow HCO_3^-(aq)$. The solution now contains 0.0200 mol of HCO_3^- ions; no gas is produced, because each carbonate ion has accepted only its first proton. (b) $HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$. (c) Moles of CO_2 = 0.0200 mol. Volume at STP = 0.0200 × 22.4 = 0.448 L. (d) $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$.

B5. (a) 0.0150 mol. HCl is a strong acid and ionises completely: every dissolved HCl molecule has donated its proton to a water molecule, so the moles of H_3O^+ equal the moles of acid dissolved. (b) The net ionic equation $H^+ + OH^- \rightarrow H_2O$ runs 1 : 1, so moles of NaOH = 0.0150 mol. (c) Volume = moles ÷ concentration = 0.0150 ÷ 0.150 = 0.100 L, which is 100 mL. (d) Still 0.0150 mol. Neutralisation collects every proton the acid can donate in total: as OH^- removes H_3O^+, more CH_3COOH molecules ionise to replace it, and this continues until all 0.0150 mol of protons have been transferred. Weakness describes the small fraction ionised at any one instant, not the total proton capacity. (e) Concentration = 0.0150 ÷ 0.250 = 0.0600 mol/L. The moles of NaOH do not change: dilution spreads the same 0.0150 mol of protons through more water, and neutralisation still has to collect all of them.

B6. (a) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$. (b) Acid: H_2O, which donates the proton; base: NH_3, which accepts it. (c) NH_3/NH_4^+ and H_2O/OH^-, each pair differing by one proton. (d) Only 2 of the 8 NH_3 molecules have accepted a proton, and the rest remain intact NH_3: proton acceptance is partial and reaches an equilibrium, which is what makes ammonia a weak base. (e) No NH_3 molecules would remain: the snapshot would show 8 NH_4^+ and 8 OH^-, with all 8 water molecules in the region used up as proton donors, because a strong base accepts a proton from water completely. (f) 2 of 8 molecules, which is 2 ÷ 8 × 100 = 25%.

C1. (a) $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$. (b) HSO_4^-. (c) $HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$. The equilibrium arrow is correct because the second proton is donated only partially: HSO_4^- is a weak acid, so most HSO_4^- ions stay intact, unlike the first step, which goes to completion. (d) $HSO_4^- + H_3O^+ \rightarrow H_2SO_4 + H_2O$. HSO_4^- is amphiprotic: it donates a proton in part (c) and accepts one here. (e) A conjugate pair differs by exactly one proton. SO_4^{2-} is two protons away from H_2SO_4, so it cannot be its conjugate base; SO_4^{2-} is the conjugate base of HSO_4^-. Removing one proton from H_2SO_4 leaves HSO_4^-.

C2. (a) 0.10 mol/L. HCl is a strong acid and ionises completely, so every HCl molecule becomes an H_3O^+ ion and the hydronium concentration equals the concentration of acid dissolved. (b) 0.10 ÷ (1.3×10^{-3}) = 77 times larger (76.9 before rounding). (c) Percentage ionised = (1.3×10^{-3} ÷ 0.10) × 100 = 1.3%. (d) Neither: they need the same amount. Both acids are monoprotic and both solutions hold 0.10 mol of donatable protons per litre; neutralisation collects the total proton capacity, and the weak acid ionises further as OH^- is added, so equal volumes of equal concentration need equal moles of NaOH. (e) Hydrochloric acid: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. Acetic acid: $CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)$, the weak acid written whole. (f) The argument confuses the fraction ionised at one instant with the total number of protons the acid can donate, and neutralisation is paid from the total: as OH^- consumes H_3O^+, more CH_3COOH molecules ionise until every proton has been transferred.

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C3. (a) NH_4^+/NH_3 and H_3O^+/H_2O. (b) H_3O^+ is the strongest acid present. Proton transfer runs most readily from the stronger acid to the stronger base, and here that is the reverse direction (H_3O^+ donating to NH_3), so the forward transfer happens only to a small extent and the equilibrium lies to the left, with most nitrogen present as NH_4^+. (c) Amphiprotic. Water's role is set by its partner: against a stronger acid it accepts a proton, against a base it donates one. (d) The NH_4^+ ions donate protons to water, $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$. The equilibrium lies to the left, but the small amount of H_3O^+ produced is still more than pure water contains on its own, so the solution registers as acidic. (e) $NH_3 + H_3O^+ \rightarrow NH_4^+ + H_2O$. Acid: H_3O^+, donating the proton; base: NH_3, accepting it. This is the reverse of the part (a) equilibrium, which is why adding ammonia to an acid consumes the H_3O^+.

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