ANSWER KEY Answers and explanations appear after each question, marked by a bar at the left. Answers written in table cells are in *italics*.

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AP Chemistry Unit 4 Study Guide | ANSWER KEY

Chemical Reactions (Topics 4.1 to 4.9)

Show all work with units and correct significant figures. When a question asks you to explain, make a claim and support it with a particle-level reason.

Constants	$N_A = 6.022 \times 10^{23}$ mol^{-1}	R = 0.08206 L·atm/(mol·K)	K = °C + 273	1 atm = 760 torr
Equations	moles = mass ÷ molar mass	molarity = mol solute ÷ L solution	PV = nRT	% yield = (actual ÷ theoretical) × 100

Part 1. Physical and Chemical Changes (Topics 4.1 and 4.4)

1. For each process, decide whether it is a physical change or a chemical change, and give the evidence or particle-level reason.

Process	Physical or chemical?	Evidence or reason
Dry ice, CO_2(s), turns directly into a gas	*Physical*	*Only the attractions between CO_2 molecules are overcome. The molecules are still CO_2.*
A magnesium ribbon burns with a bright white light	*Chemical*	*Light and heat are released, and a new substance (MgO) forms.*
Mixing clear NaCl(aq) and $AgNO_3$(aq) forms a white solid	*Chemical*	*A precipitate (AgCl) forms. Ag^+ and Cl^- combine into a new compound.*
Sugar dissolves in water	*Physical*	*Sugar molecules separate and are surrounded by water. No covalent bonds in the sugar break.*
An iron nail slowly turns into orange-brown rust	*Chemical*	*Color change and a new substance (iron oxide) with a different composition.*

2. A student claims, "Boiling water is a chemical change because bubbles of gas form." Do you agree or disagree? Explain in terms of the bonds or forces that are overcome.

> Disagree. The bubbles are water vapor, H_2O(g), not a new substance. Boiling overcomes the hydrogen bonds between H_2O molecules (intermolecular forces). The covalent O−H bonds inside each molecule do not break, so the composition does not change. Formation of a gas is only evidence of a chemical change when the gas is a new substance.

3. Dissolving NaCl(s) in water can be argued to be either a physical or a chemical process.

(a) Give one argument that it is a physical process.

> The water can be evaporated to recover the same NaCl(s). No new substance forms, and the ions are still Na^+ and Cl^-.

(b) Give one argument that it is a chemical process.

> The ionic bonds (attractions between Na^+ and Cl^- in the lattice) are broken, and new ion-dipole interactions form between the ions and water molecules. Breaking and forming interactions of this strength is usually associated with chemical processes.

Part 2. Net Ionic Equations (Topic 4.2)

4. Clear solutions of lead(II) nitrate, $Pb(NO_3)_2$(aq), and potassium iodide, KI(aq), are mixed, and a bright yellow solid forms. (All potassium and nitrate salts are soluble.)

(a) Write the balanced molecular equation, with states.

> $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$. KNO_3 is soluble, so the yellow solid must be PbI_2.

(b) Write the complete ionic equation.

$Pb^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow PbI_2(s) + 2K^+(aq) + 2NO_3^-(aq)$

(c) Write the net ionic equation and name the spectator ions.

$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$. Spectator ions: K^+ and NO_3^-. They appear unchanged on both sides, so they are removed.

5. Write the net ionic equation for each reaction. Remember that weak acids and solids are written as intact formulas, not as ions.

(a) $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. HCl is a strong acid and NaOH is a strong base, so both are written as ions; Na^+ and Cl^- are spectators.

(b) $HC_2H_3O_2(aq) + NaOH(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l)$ (acetic acid is a weak acid)

$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$. Acetic acid is mostly un-ionized in water, so it stays as a molecule. Na^+ is the only spectator.

(c) $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$. The solid stays intact; Cl^- is the spectator.

6. Physical changes can be shown with balanced equations too. Write a balanced equation, with states, for each change.

(a) Solid iodine sublimes.

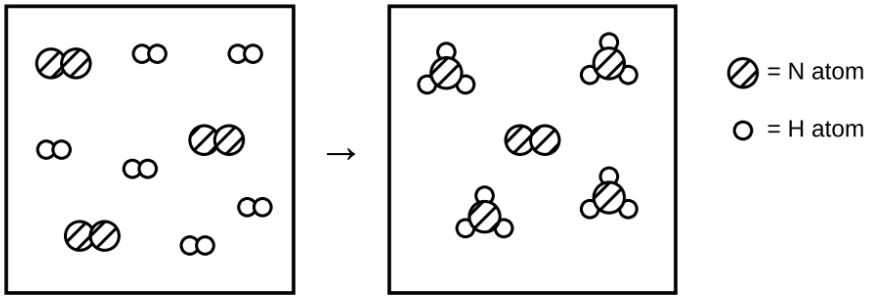
$I_2(s) \rightarrow I_2(g)$

(b) Solid ammonium nitrate dissolves in water.

$NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$. The polyatomic ions stay intact.

Part 3. Representations of Reactions (Topic 4.3)

7. Nitrogen and hydrogen react to form ammonia: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. The left box shows a reaction mixture before any reaction occurs.



(a) In the right box, draw the particles present after the reaction goes to completion.

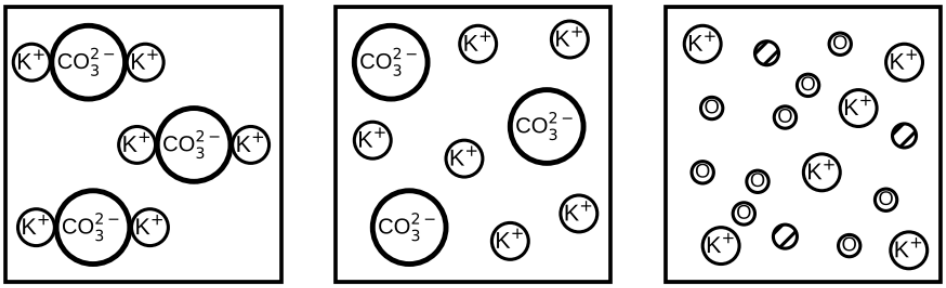
4 NH_3 molecules and 1 leftover N_2 molecule (shown in the key diagram). The 6 H_2 react with 2 N_2 to make 4 NH_3. Atoms are conserved: 6 N and 12 H before and after.

(b) Identify the limiting reactant. Justify your answer using the diagram.

H_2 is limiting. 3 N_2 would need 9 H_2, but only 6 H_2 are present. All the H_2 is used up while 1 N_2 is left over.

8. Which box best represents $K_2CO_3(aq)$? Circle one: **A** **B** **C**. Then explain what is wrong with each of the other two boxes.

Water molecules are not shown. Hatched small circles in C are carbon atoms.



> **B.** K_2CO_3 is an ionic compound, so in water it separates into K^+ and CO_3^{2-} ions in a 2 to 1 ratio.
> A is wrong because the ions are still paired as formula units; dissolving separates them. C is wrong because carbonate is a polyatomic ion. Its covalent C−O bonds do not break when it dissolves.

Part 4. Stoichiometry (Topic 4.5)

9. Aluminum reacts with chlorine gas: $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$. A student reacts 10.0 g of Al with 35.0 g of Cl_2. (Molar masses: Al 26.98, Cl_2 70.90, $AlCl_3$ 133.33 g/mol)

(a) Identify the limiting reactant. Show your work.

> mol Al = 10.0 g ÷ 26.98 g/mol = 0.3706 mol. mol Cl_2 = 35.0 g ÷ 70.90 g/mol = 0.4937 mol.
> Al would need 0.3706 × (3/2) = 0.5560 mol Cl_2, but only 0.4937 mol is present, so **Cl_2 is limiting**.

(b) Calculate the theoretical yield of $AlCl_3$ in grams.

> 0.4937 mol Cl_2 × (2 mol $AlCl_3$ / 3 mol Cl_2) × 133.33 g/mol = **43.9 g $AlCl_3$**

(c) What mass of the excess reactant remains?

> Al used = 0.4937 mol Cl_2 × (2 mol Al / 3 mol Cl_2) = 0.3291 mol Al × 26.98 g/mol = 8.88 g. Remaining = 10.0 g − 8.88 g = **1.1 g Al** (subtraction keeps the tenths place)

(d) The student actually collects 38.2 g of $AlCl_3$. Calculate the percent yield.

> % yield = (38.2 g ÷ 43.88 g) × 100 = **87.1%** (using the unrounded theoretical yield)

10. If the student had used 20.0 g of Al with the same 35.0 g of Cl_2, the mass of $AlCl_3$ formed would (circle one) **increase** **decrease** **stay the same**. Explain.

> **Stay the same.** Cl_2 is still the limiting reactant, so it still makes only 43.9 g of $AlCl_3$. Adding more of the excess reactant just leaves more Al unreacted.

11. Potassium chlorate decomposes when heated: $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$. What volume of O_2 forms at 1.00 atm and 298 K when 4.90 g of $KClO_3$ (122.55 g/mol) decomposes completely?

> mol $KClO_3$ = 4.90 g ÷ 122.55 g/mol = 0.03998 mol. mol O_2 = 0.03998 × (3/2) = 0.05998 mol
> $V = nRT/P$ = (0.05998 mol)(0.08206 L·atm/(mol·K))(298 K) ÷ 1.00 atm = **1.47 L O_2**

12. Silver nitrate solution is added to 25.0 mL of 0.150 M $CaCl_2(aq)$. The net ionic equation is $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.

(a) What volume of 0.250 M $AgNO_3$ is needed to precipitate all of the chloride ions?

> mol Cl^- = 0.150 mol/L × 0.0250 L × 2 = 0.00750 mol (each $CaCl_2$ gives 2 Cl^-). V = 0.00750 mol ÷ 0.250 mol/L = 0.0300 L = **30.0 mL**

(b) What mass of AgCl (143.32 g/mol) forms?

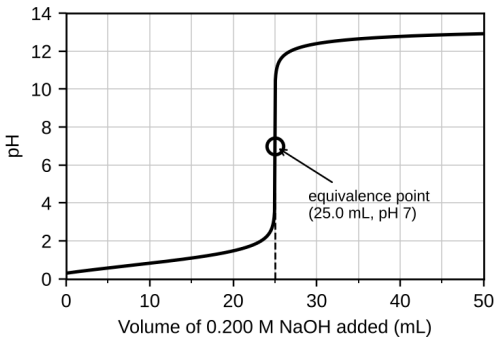
> 0.00750 mol AgCl × 143.32 g/mol = **1.07 g AgCl**

Part 5. Introduction to Titration (Topic 4.6)

13. A 20.0 mL sample of HCl(aq) is titrated with 0.100 M NaOH. The indicator changes color when 24.6 mL of NaOH has been added. Calculate the concentration of the HCl.

> mol NaOH = 0.100 mol/L × 0.0246 L = 0.00246 mol. HCl and NaOH react 1 to 1, so mol HCl = 0.00246 mol. [HCl] = 0.00246 mol ÷ 0.0200 L = **0.123 M**

14. The graph shows the titration of a 10.0 mL sample of HCl(aq) with 0.200 M NaOH.



(a) Circle the equivalence point on the graph and estimate the volume of NaOH added there.

> About 25.0 mL, at the center of the steep vertical jump in pH (pH 7 for a strong acid with a strong base).

(b) Calculate the concentration of the HCl.

> mol NaOH = 0.200 mol/L × 0.0250 L = 0.00500 mol = mol HCl. [HCl] = 0.00500 mol ÷ 0.0100 L = **0.500 M**

(c) Explain the difference between the equivalence point and the endpoint.

> The equivalence point is when the moles of NaOH added exactly equal the moles of HCl, so the analyte is totally consumed. The endpoint is the observed color change of the indicator. A good indicator changes color very close to the equivalence point.

15. Before a titration, a student rinses the buret with distilled water but not with the NaOH solution, so a few drops of water dilute the NaOH. Will the calculated concentration of the acid be too high, too low, or unaffected? Explain.

> **Too high.** The diluted NaOH is weaker than its labeled concentration, so a larger volume is needed to reach the endpoint. The student multiplies that larger volume by the labeled molarity, which overestimates the moles of NaOH, and therefore the moles and concentration of acid.

Part 6. Types of Chemical Reactions (Topic 4.7)

16. Classify each reaction as **precipitation**, **acid-base**, or **redox**, and give the evidence.

Reaction	Type	Evidence
$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$	*Precipitation*	*Ions combine to form an insoluble solid. No oxidation numbers change.*
$NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$	*Acid-base*	*An H^+ is transferred from HCl to NH_3, forming NH_4^+.*
$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$	*Redox*	*Zn goes from 0 to +2 (oxidized); Cu goes from +2 to 0 (reduced).*
$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$	*Redox (combustion)*	*C goes from −4 to +4 (oxidized); O goes from 0 to −2 (reduced).*

17. Write the balanced equation for the complete combustion of propane, $C_3H_8(g)$. Then explain why combustion is classified as a redox reaction.

> $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$. Combustion is redox because carbon is oxidized (average oxidation number −8/3 to +4) and oxygen is reduced (0 to −2). Electrons are transferred from the fuel to O_2.

18. Using only the rule that all sodium, potassium, ammonium, and nitrate salts are soluble, circle every compound that *must* be soluble in water:

KBr NH_4Cl $BaSO_4$ $Pb(NO_3)_2$ AgCl Na_3PO_4

> KBr, NH_4Cl, $Pb(NO_3)_2$, and Na_3PO_4. Each contains K^+, NH_4^+, NO_3^-, or Na^+. The rule says nothing about $BaSO_4$ or AgCl (both are actually insoluble, which is why they show up as precipitates).

19. Write **T** or **F** for each statement. Rewrite each false statement so it is true.

Statement	T or F	Correction (if false)
The complete combustion of a hydrocarbon is an acid-base reaction.	*F*	*Combustion is a **redox** reaction: the fuel is oxidized and O_2 is reduced.*
In a precipitation reaction, the oxidation numbers of the ions change.	*F*	*In a precipitation reaction the oxidation numbers do **not** change; ions just combine into an insoluble solid.*
In a redox reaction, electrons move from the species oxidized to the species reduced.	*T*	

Part 7. Acid-Base Reactions (Topic 4.8)

20. For each reaction, identify the Brønsted-Lowry acid, base, conjugate acid, and conjugate base.

Reaction	Acid	Base	Conj. acid	Conj. base
$HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)$	*HF*	*H_2O*	*H_3O^+*	*F^-*
$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$	*H_2O*	*NH_3*	*NH_4^+*	*OH^-*
$HSO_4^-(aq) + CO_3^{2-}(aq) \rightleftharpoons SO_4^{2-}(aq) + HCO_3^-(aq)$	*HSO_4^-*	*CO_3^{2-}*	*HCO_3^-*	*SO_4^{2-}*

> The acid is the species that donates H^+; its conjugate base is what remains. The base accepts H^+; its conjugate acid is the species that forms.

21. Use the first two reactions in the table above to explain why water is described as able to act as both an acid and a base.

> With HF, water accepts a proton (acts as a base) to form H_3O^+. With NH_3, water donates a proton (acts as an acid) and becomes OH^-. Its structure lets it gain or lose H^+ depending on what it reacts with.

Part 8. Oxidation-Reduction (Redox) Reactions (Topics 4.7 and 4.9)

22. Assign the oxidation number of the named atom in each species.

Cr in $Cr_2O_7^{2-}$	S in SO_4^{2-}	N in NH_4^+	Mn in MnO_4^-	O in H_2O_2
+6	*+6*	*−3*	*+7*	*−1*

> Use O = −2 (except in peroxides like H_2O_2, where O = −1) and H = +1, and make the sum equal the charge. Example: 2Cr + 7(−2) = −2, so Cr = +6.

23. Consider $2Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{4+}(aq)$.

(a) Which species is oxidized and which is reduced? Explain.

> Sn^{2+} is oxidized (+2 to +4, loses electrons). Fe^{3+} is reduced (+3 to +2, gains electrons).

(b) Write the oxidation and reduction half-reactions. How many electrons are transferred in the balanced reaction?

> Oxidation: $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$. Reduction: $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (multiplied by 2). **2 electrons** are transferred from Sn^{2+} to the two Fe^{3+} ions.

24. Use half-reactions to balance this reaction in acidic solution: $MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq)$

> Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (add H_2O to balance O, H^+ to balance H, then e^- to balance charge).
> Oxidation: $Fe^{2+} \rightarrow Fe^{3+} + e^-$, multiplied by 5 so the electrons cancel.
> Overall: **$MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$**. Check: charge is +17 on each side.

25. Using the balanced equation from the previous question, a 25.00 mL sample of $Fe^{2+}(aq)$ requires 18.40 mL of 0.0200 M $KMnO_4$ to reach the endpoint (the first lasting pink color). Calculate $[Fe^{2+}]$.

> mol MnO_4^- = 0.0200 mol/L × 0.01840 L = 3.68×10^{-4} mol. mol Fe^{2+} = 5 × 3.68×10^{-4} = 1.84×10^{-3} mol. $[Fe^{2+}]$ = 1.84×10^{-3} mol ÷ 0.02500 L = **0.0736 M**

26. A student writes $Cu(s) + Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + Ag(s)$ and says it is balanced because the atoms balance. Explain the error and write the correctly balanced equation.

> Charge is not balanced: the left side has a total charge of +1 and the right side +2. Cu loses 2 electrons but each Ag^{+} gains only 1, so 2 Ag^{+} are needed: **$Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$**.