AL 化学逆袭宝典

CIE AS Chemistry • Basis

新灯塔国际教育

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1 Atomic Structure（原子结构）

Protons, neutrons and electrons; isotopes; electron configurations (H–Kr basics); simple ionisation energy ideas.

Rules 规则

- **Atomic number** Z（原子序数）: number of protons in the nucleus; defines the element.
- **Mass number** A（质量数）: protons + neutrons; $A = Z + n$; neutrons $n = A - Z$.
- **Neutral atom**（中性原子）: number of electrons $= Z$.
- **Ions**（离子）: cation X^{n+}: electrons $= Z - n$; anion X^{n-}: electrons $= Z + n$.
  e.g. Na ($Z = 11$) $\rightarrow Na^+$ has 10 e; O ($Z = 8$) $\rightarrow O^{2-}$ has 10 e.
- **Isotopes**（同位素）: same Z, different A (different neutron number); same chemical properties.
- **Relative atomic mass** A_r（相对原子质量）: weighted mean of isotopic masses using percentage abundances:
  $A_r = \dfrac{\sum(A_i \times a_i)}{100}$, where a_i is the percentage abundance of isotope i.
- **Electron shells**（电子层）: principal shells $n = 1, 2, 3, \ldots$; capacity roughly $2n^2$ (AS level).
- **Subshells & filling**（亚层与填充）: s (2 e), p (6 e), d (10 e); fill in order $1s\,2s\,2p\,3s\,3p\,4s\,3d\,4p$.
  Common AS exceptions: Cr $4s^1\,3d^5$; Cu $4s^1\,3d^{10}$.
  For $3d$-block cations, remove $4s$ electrons before $3d$ (e.g. Fe: $4s^2\,3d^6 \rightarrow Fe^{2+}$: $3d^6$).
- **First / second ionisation energy**（第一/第二电离能）: energy to remove 1 mol of electrons from 1 mol of gaseous atoms / ions:
  1st: $X(g) \longrightarrow X^+(g) + e^-$
  2nd: $X^+(g) \longrightarrow X^{2+}(g) + e^-$
  (always gaseous species; 2nd IE is larger than 1st IE because the electron is removed from a positive ion).
- **Successive IE & group**（连续电离能与族）: a large jump after the nth IE means n outer-shell electrons (e.g. jump after $IE_2 \Rightarrow$ Group 2).
- **IE trends**（电离能趋势）: generally **increases** across a period (nuclear charge ↑, similar shielding); generally **decreases** down a group (atomic radius ↑, more shielding).
- **Period 2 / 3 anomalies**（周期异常，简要）: Be > B and Mg > Al (full s^2 harder than removing a p electron);
  N > O and P > S (half-filled p^3 more stable; paired p electron easier to remove).

Worked Examples 例题

Example 1. An atom of chlorine has $Z = 17$ and $A = 35$. State the number of protons, neutrons and electrons in a neutral $^{35}\mathrm{Cl}$ atom.

Protons $= 17$; neutrons $= 35 - 17 = 18$; electrons $= 17$.

Example 2. Write the full electron configuration of O ($Z = 8$) and of O^{2-}.

O: $1s^2\,2s^2\,2p^4$.

O^{2-}: $1s^2\,2s^2\,2p^6$.

Example 3. Define first ionisation energy. Write the equation for the first ionisation of sodium.

Energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous unipositive ions.

$\mathrm{Na(g)} \longrightarrow \mathrm{Na^+(g)} + \mathrm{e^-}$.

Basic Drill （入门综合）· 32

1. Define atomic number, Z.

2. Define mass number, A.

3. State the charge and relative mass of a proton.

4. State the charge and relative mass of a neutron.

5. State the charge and relative mass of an electron.

6. For ^{12}C, state protons, neutrons and electrons (neutral atom).

7. For ^{23}Na, state protons, neutrons and electrons (neutral atom).

8. For ^{16}O, state protons, neutrons and electrons (neutral atom).

9. For ^{35}Cl, state protons, neutrons and electrons (neutral atom).

10. For ^{40}Ca, state protons, neutrons and electrons (neutral atom).

11. How many electrons are in Na^+?

12. How many electrons are in O^{2-}?

13. How many electrons are in Cl^{-}?

14. How many electrons are in Al^{3+}?

15. Define isotopes.

16. ^{35}Cl and ^{37}Cl are isotopes. What is the same and what differs?

17. Write the electron configuration of H ($Z = 1$).

18. Write the electron configuration of He ($Z = 2$).

19. Write the electron configuration of C ($Z = 6$).

20. Write the electron configuration of N ($Z = 7$).

21. Write the electron configuration of F ($Z = 9$).

22. Write the electron configuration of Ne ($Z = 10$).

23. Write the electron configuration of Na ($Z = 11$).

24. Write the electron configuration of Mg ($Z = 12$).

25. Write the electron configuration of Al ($Z = 13$).

26. Write the electron configuration of Cl ($Z = 17$).

27. Write the electron configuration of Ar ($Z = 18$).

28. Write the electron configuration of K ($Z = 19$).

29. Write the electron configuration of Ca ($Z = 20$).

30. Define first ionisation energy.

31. Write the equation for the first ionisation of magnesium.

32. Across Period 2 from Li to Ne, does first IE generally increase or decrease?

Practice （综合练习）· 14

Practice

1. An ion X^{2+} has 18 electrons and mass number 40. Deduce Z, A and the number of neutrons; identify X.

2. Write the electron configurations of S and S^{2-}.

3. Write the electron configurations of Fe ($Z = 26$) and Fe^{2+} (remove $4s$ electrons first).

4. Write the electron configuration of Cu ($Z = 29$) using the common AS exception ($4s^1\,3d^{10}$).

5. $^{39}K^{+}$ and $^{40}Ca^{2+}$ are isoelectronic. State what this means and confirm by counting electrons.

6. Calculate the relative atomic mass of chlorine given ^{35}Cl (75%) and ^{37}Cl (25%).

7. An element has isotopes of mass 24 (80%) and 25 (20%). Calculate A_r.

8. State which has the higher first IE: Na or Mg. Give a one-line reason.

9. State which has the higher first IE: Na or K. Give a one-line reason.

10. Write equations for the first and second ionisation of calcium.

11. Explain why the second IE of Mg is larger than its first IE.

12. Give the full electron configuration of Br ($Z = 35$).

13. Give the full electron configuration of Kr ($Z = 36$).

14. An atom has configuration $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$. Identify the element and state its group.

Further Practice （进阶）· 10

Further Practice

A. Particles and isotopes （真题常考 · 基础计数与符号）

1. Deduce the numbers of protons, neutrons and electrons in $^{56}\mathrm{Fe}^{3+}$.

2. A species has 17 protons, 18 neutrons and 18 electrons. Give its full isotopic symbol (including charge), e.g. $^{A}\mathrm{X}^{n\pm}$.

3. State what is meant by isoelectronic. Confirm that $^{39}\mathrm{K}^{+}$ and $^{40}\mathrm{Ca}^{2+}$ are isoelectronic by counting electrons.

B. Relative atomic mass （丰度加权计算）

4. Two isotopes of an element have $A = 63$ (70%) and $A = 65$ (30%). Calculate A_r and suggest the element.

5. An element has two isotopes. The lighter isotope ($A = 10$) has abundance 19.9%; the heavier isotope has $A = 11$. Calculate A_r to 3 s.f. and identify the element.

C. Electron configurations （书写排布与常见例外）

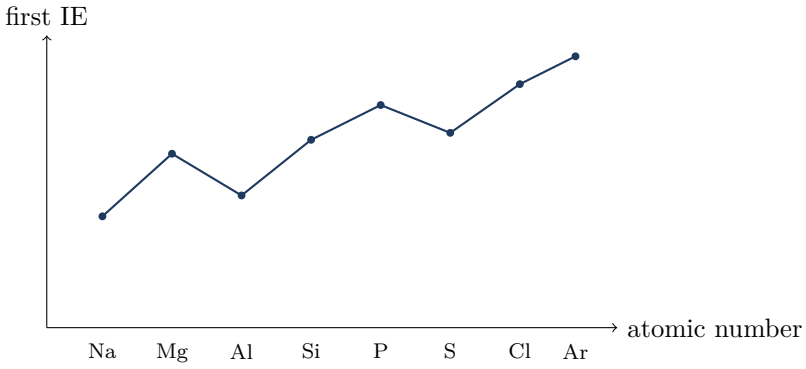
6. Write the electron configuration of Cr ($Z = 24$) using the common AS exception ($4s^1\,3d^5$).

7. Write the full electron configurations of Ni ($Z = 28$) and of Ni^{2+} (remove $4s$ electrons before $3d$).

8. Write the electron configuration of Cu ($Z = 29$) using the common AS exception ($4s^1\,3d^{10}$), and give the configurations of Cu^+ and Cu^{2+}.

9. An atom has configuration $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$. Identify the element and state its group.

D. First IE graph (Period 3) （第一电离能图像）



10. The sketch shows the *trend* in first ionisation energy across Period 3. (i) Describe the overall trend from Na to Ar.
    (ii) Explain the drop from Mg to Al.
    (iii) Explain the drop from P to S.

Advanced Practice （培优）· 10

Advanced Practice

A. Species deduction from combined data （IE 跃迁 + 粒子数双约束）

1. Successive IE of element E show a very large jump after IE_2. A monatomic ion of E has charge +2, contains 10 electrons, and has mass number 24. Deduce Z, the number of neutrons in this ion, and identify E.

2. An ion Y^{n-} is isoelectronic with Ar and contains 18 neutrons. Its mass number is 34. Deduce Z, n and identify Y.

B. Mass spectra of isotopes （丰度比 → 峰强比）

3. Bromine has two isotopes, ^{79}Br and ^{81}Br, of roughly equal abundance. For BrCl, predict how many peaks of *different* m/z appear near M^+ and identify which m/z is the most intense (use $^{35}Cl/^{37}Cl \approx 3:1$).

4. Chlorine isotopes $^{35}Cl{:}^{37}Cl \approx 3:1$. Explain why CCl_4 shows a cluster of molecular-ion peaks, and state which chlorine isotopic combination gives the lowest m/z peak.

C. Successive ionisation energies （定族 + 区分内外层）

5. The first five IE of R ($kJ\,mol^{-1}$) are 786, 1580, 3230, 4360, 16100. Deduce the group of R. Explain why IE_5 is so much larger than IE_4.

6. Na and Mg are consecutive elements. Explain why the *second* IE of Na is much larger than the second IE of Mg, even though Na has the smaller nuclear charge.

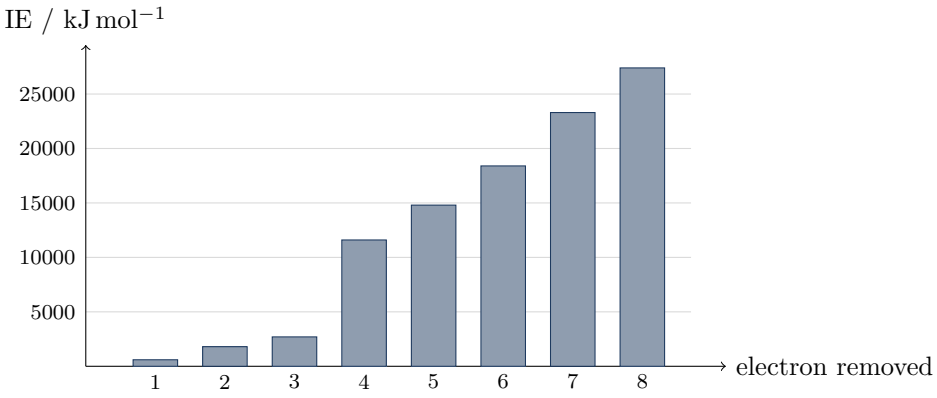
7. Write equations for the first and second ionisation of aluminium, and state which electron (subshell) is removed in each step.

D. IE trends and anomalies （必须同时给趋势机制 + 异常机制）

8. First IE drops from Mg to Al and from P to S. Explain *each* drop using electron configurations (do not only quote "anomaly").

9. Across Period 3, first IE of Ar is the highest. Explain why, referring to nuclear charge, shielding and the full outer shell.

E. Successive IE graph （连续电离能柱状图）



10. The bar chart shows successive ionisation energies of element Q. (i) Deduce the group of Q.
(ii) Explain the large jump between IE_3 and IE_4.
(iii) Write the equation for the third ionisation of Q (use Q).

2 Atoms, Molecules & Stoichiometry（原子、分子与化学计量）

Relative masses, the mole, formulas & equations; reacting masses, gas volumes, solutions, limiting reagent and percentage yield.

Data unless stated: H 1.0, C 12.0, N 14.0, O 16.0, Na 23.0, Mg 24.0, Al 27.0, S 32.0, Cl 35.5, K 39.0, Ca 40.0, Fe 56.0, Cu 64.0, Zn 65.0, Ag 108, Ba 137. Molar gas volume at r.t.p. = 24.0 $\mathrm{dm^3\,mol^{-1}}$.

Rules 规则

- **Unified atomic mass unit**（统一原子质量单位）：1 u is $\frac{1}{12}$ of the mass of one ^{12}C atom.
- **Relative atomic / molecular / formula mass**（相对原子/分子/式量）：A_r, M_r are ratios to 1 u（无单位）.
- **Mole & Avogadro constant**（摩尔与阿伏加德罗常数）：1 mol contains $L = 6.02 \times 10^{23}$ specified entities.
- **Amount of substance**（物质的量）：$n = \dfrac{m}{M}$, $n = \dfrac{V_\mathrm{gas}}{V_\mathrm{m}}$, $c = \dfrac{n}{V_\mathrm{sol}}$, $n = cV$. （M in $\mathrm{g\,mol^{-1}}$; V_sol in $\mathrm{dm^3}$.）
- **Empirical vs molecular formula**（经验式与分子式）：empirical = simplest whole-number ratio; molecular = actual numbers; $M_\mathrm{r}(\text{molecular}) = k \times M_\mathrm{r}(\text{empirical})$.
- **Hydrates**（水合物）：e.g. $\mathrm{CuSO_4 \cdot \mathit{x}\,H_2O}$; water of crystallisation from mass loss on heating: $\mathrm{CuSO_4 \cdot 5\,H_2O\,(s) \longrightarrow CuSO_4(s) + 5\,H_2O\,(g)}$.
- **Equations**（方程式）：balance atoms & charge; ionic equations omit spectator ions; include state symbols.
  e.g. full: $\mathrm{AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)}$;
  ionic: $\mathrm{Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)}$.
- **Limiting reagent**（限量试剂）：the reactant that runs out first controls product amount.
  e.g. $\mathrm{2\,H_2 + O_2 \longrightarrow 2\,H_2O}$: if $n(\mathrm{H_2})/2 < n(\mathrm{O_2})$, then $\mathrm{H_2}$ is limiting.
- **Percentage yield**（百分产率）：$\%\,\text{yield} = \dfrac{\text{actual mass}}{\text{theoretical mass}} \times 100\%$.
- **Significant figures**（有效数字）：match data given / asked; quote units where required.

Worked Examples 例题

Example 1. Calculate M_r of $\mathrm{H_2O}$ and the amount of substance in 9.00 g of water.

$M_r = 2 \times 1.0 + 16.0 = 18.0.$

$n = \dfrac{9.00}{18.0} = 0.500\,\text{mol}.$

Example 2. A compound contains 40.0% C, 6.7% H and 53.3% O by mass. Find its empirical formula.

Assume 100 g: $n(\text{C}) = \dfrac{40.0}{12.0} = 3.33, \quad n(\text{H}) = \dfrac{6.7}{1.0} = 6.7, \quad n(\text{O}) = \dfrac{53.3}{16.0} = 3.33.$

Divide by 3.33: ratio $C : H : O = 1 : 2 : 1 \Rightarrow$ empirical formula CH_2O.

Example 3. Find n in 25.0 cm^3 of 0.200 mol dm^{-3} HCl. This acid reacts completely with NaOH: $\text{HCl} + \text{NaOH} \longrightarrow \text{NaCl} + \text{H}_2\text{O}$. Find the volume of 0.100 mol dm^{-3} NaOH required.

$n(\text{HCl}) = 0.200 \times 0.0250 = 0.00500\,\text{mol}.$

Mole ratio $1 : 1 \Rightarrow n(\text{NaOH}) = 0.00500\,\text{mol}.$

$V(\text{NaOH}) = \dfrac{0.00500}{0.100} = 0.0500\,\text{dm}^3 = 50.0\,\text{cm}^3.$

Example 4. 2.40 g of Mg burns in excess oxygen. Calculate the mass of MgO formed.

$$2\,\text{Mg} + \text{O}_2 \longrightarrow 2\,\text{MgO}$$

$n(\text{Mg}) = \dfrac{2.40}{24.0} = 0.100\,\text{mol}.$

From the equation, $n(\text{MgO}) = n(\text{Mg}) = 0.100\,\text{mol}.$

$m(\text{MgO}) = 0.100 \times 40.0 = 4.00\,\text{g}.$

Basic Drill （入门综合）· 40

1. Define the unified atomic mass unit.

2. Define relative atomic mass, A_r.

3. Define relative molecular mass, M_r.

4. State what is meant by one mole of entities.

5. Calculate A_r of chlorine given isotopes $^{35}\mathrm{Cl}$ (75.0%) and $^{37}\mathrm{Cl}$ (25.0%).

6. Calculate M_r of $\mathrm{H_2O}$.

7. Calculate M_r of $\mathrm{Ca(OH)_2}$.

8. Find n in 8.80 g of $\mathrm{CO_2}$.

9. Find the mass of 0.200 mol of $\mathrm{Na_2CO_3}$.

10. How many molecules are in 0.0100 mol of $\mathrm{H_2O}$? (Use $L = 6.02 \times 10^{23}\,\mathrm{mol^{-1}}$.)

11. Write the formula of sodium chloride.

12. Write the formula of magnesium oxide.

13. Write the formula of potassium sulfate.

14. Write the formula of calcium carbonate.

15. Balance: $\mathrm{H_2 + O_2 \longrightarrow H_2O}$.

16. Balance: $\mathrm{Fe + O_2 \longrightarrow Fe_2O_3}$.

17. Balance and add state symbols: $Na_2CO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$.

18. Write an ionic equation for $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$.

19. A compound contains 40.0% C, 6.7% H and 53.3% O by mass. Find its empirical formula.

20. Empirical formula CH_2O; $M_r = 60.0$. Find the molecular formula.

21. Empirical formula NO_2; $M_r = 92.0$. Find the molecular formula.

22. Calculate the percentage by mass of oxygen in H_2O.

23. Calculate the percentage by mass of water in $CuSO_4 \cdot 5\,H_2O$.

24. What volume at r.t.p. is occupied by 0.250 mol of O_2?

25. What amount of gas occupies $4.80\,dm^3$ at r.t.p.?

26. Burning 0.100 mol of CH_4 completely: find the volume of CO_2 formed at r.t.p.
$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$.

27. Find n of solute in $250\,cm^3$ of $0.200\,mol\,dm^{-3}$ HCl.

28. What mass of NaOH is needed to prepare $500\,\text{cm}^3$ of $0.100\,\text{mol}\,\text{dm}^{-3}$ solution?

29. Find the concentration (in $\text{mol}\,\text{dm}^{-3}$) if $5.85\,\text{g}$ of NaCl is dissolved to make $250\,\text{cm}^3$ of solution.

30. $25.0\,\text{cm}^3$ of $0.100\,\text{mol}\,\text{dm}^{-3}$ HCl is neutralised by NaOH. Find $n(\text{HCl})$.

31. Using Q30 and $\text{HCl} + \text{NaOH} \longrightarrow \text{NaCl} + \text{H}_2\text{O}$, find $n(\text{NaOH})$ required.

32. Find the volume of $0.100\,\text{mol}\,\text{dm}^{-3}$ NaOH needed in Q31.

33. From $2\,\text{H}_2 + \text{O}_2 \longrightarrow 2\,\text{H}_2\text{O}$, what mass of water forms from $2.00\,\text{g}$ of H_2 (excess O_2)?

34. What mass of CaCO_3 is needed to produce $2.40\,\text{dm}^3$ of CO_2 at r.t.p.?
$\text{CaCO}_3 \longrightarrow \text{CaO} + \text{CO}_2$.

35. $6.50\,\text{g}$ of Zn reacts with excess HCl. Find the volume of H_2 at r.t.p.
$\text{Zn} + 2\,\text{HCl} \longrightarrow \text{ZnCl}_2 + \text{H}_2$.

36. Identify the limiting reagent when $0.100\,\text{mol}$ N_2 meets $0.200\,\text{mol}$ H_2:
$\text{N}_2 + 3\,\text{H}_2 \longrightarrow 2\,\text{NH}_3$.

37. Using Q36, find the maximum $n(\text{NH}_3)$ formed.

38. Theoretical mass of product is $10.0\,\text{g}$; actual mass obtained is $8.00\,\text{g}$. Calculate percentage yield.

39. A pure sample of $CaCO_3$ of mass 5.00 g is heated completely. Calculate the mass of CO_2 produced.
$CaCO_3 \longrightarrow CaO + CO_2$.

40. 2.40 g of Mg is burned in excess oxygen. Calculate the mass of MgO formed.
$2\,Mg + O_2 \longrightarrow 2\,MgO$.

Practice （综合练习）· 18

Practice

1. Calculate M_r of $Al_2(SO_4)_3$, then find n in 3.42 g of this salt.

2. A compound contains 80.0% C and 20.0% H by mass. $M_r = 30.0$. Find the empirical and molecular formulas.

3. 2.50 g of hydrated copper(II) sulfate $CuSO_4 \cdot x\,H_2O$ is heated to constant mass. The anhydrous residue has mass 1.60 g. Determine x.

4. Write a balanced ionic equation (with state symbols) for the reaction between aqueous silver nitrate and aqueous sodium chloride.

5. 25.0 cm^3 of 0.100 $mol\,dm^{-3}$ HCl reacts with 0.100 $mol\,dm^{-3}$ NaOH:
$HCl + NaOH \longrightarrow NaCl + H_2O$.
Find the volume of NaOH solution required for complete reaction.

6. 25.0 cm^3 of 0.200 $mol\,dm^{-3}$ H_2SO_4 reacts with NaOH:
$H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + 2\,H_2O$.
Find $n(H_2SO_4)$ and $n(NaOH)$ required.

7. 8.00 g of Fe_2O_3 is reduced by excess CO. Calculate the mass of Fe formed.

$Fe_2O_3 + 3\,CO \longrightarrow 2\,Fe + 3\,CO_2$.

8. In the reaction of Q7, what volume of CO_2 (r.t.p.) is produced from 8.00 g of Fe_2O_3?

9. A solution is made by dissolving 5.30 g of anhydrous Na_2CO_3 and making up to 250 cm^3. Calculate its concentration in mol dm^{-3}.

10. 20.0 cm^3 of the solution in Q9 reacts with HCl:
$Na_2CO_3 + 2\,HCl \longrightarrow 2\,NaCl + H_2O + CO_2$.
Find $n(Na_2CO_3)$ used and $n(HCl)$ required.

11. 6.50 g of Zn is mixed with 4.00 g of S and heated: $Zn + S \longrightarrow ZnS$. Determine the limiting reagent and the mass of ZnS formed.

12. Starting from 14.0 g of N_2 and 3.00 g of H_2, decide if a limiting reagent exists (or the amounts are stoichiometric) and find the maximum mass of NH_3.
$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$.

13. In an experiment, 11.9 g of NH_3 was actually obtained from the mixture in Q12. Calculate the percentage yield.

14. A metal M forms a chloride MCl_2. 1.00 g of M reacts with chlorine to give 2.78 g of MCl_2. Deduce A_r of M and identify M from the data list.

15. 25.0 cm^3 of ethanoic acid solution requires 20.0 cm^3 of 0.100 mol dm^{-3} NaOH.
$CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$.
Find the concentration of the ethanoic acid.

16. Heating 10.0 g of impure $CaCO_3$ with excess acid produces 2.16 dm^3 of CO_2 at r.t.p. Calculate the percentage purity.

17. In acidified manganate(VII) titration, 25.0 cm^3 of 0.100 mol dm^{-3} Fe^{2+} needs 10.0 cm^3 of 0.0500 mol dm^{-3} MnO_4^-. State the mole ratio $n(Fe^{2+}) : n(MnO_4^-)$.

18. Propane burns: $C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O$.
A mixture contains 0.0200 mol propane and 0.150 mol oxygen. Find the limiting reagent and the amount of CO_2 formed.

Further Practice （进阶）· 9

Further Practice

A. Amount of substance and simple mole calculations （$n = m/M$、气体体积基础）

1. Calculate M_r of $CaCO_3$ and the amount of substance in 10.0 g of $CaCO_3$.

2. Find the mass of 0.250 mol of Na_2CO_3.

3. What volume (at r.t.p.) is occupied by 0.150 mol of CO_2?

B. Empirical and molecular formulas; hydrates （百分组成 · 水合物）

4. An oxide of nitrogen contains 30.4% N by mass. Its M_r is 92.0. Find the empirical and molecular formulas.

5. Hydrated iron(II) sulfate $FeSO_4 \cdot x\,H_2O$ is analysed: 13.9 g of the hydrate contains 6.30 g of water of crystallisation. Find x.

6. A hydrocarbon contains 85.7% C and 14.3% H by mass; $M_r = 56$. Find the empirical and

molecular formulas.

C. Simple reacting masses and gas volumes （单步计量 · 非限量）

7. A sample of $MgCO_3$ of mass 4.20 g is heated strongly until decomposition is complete: $MgCO_3 \longrightarrow MgO + CO_2$.
   Calculate the mass of MgO formed and the volume of CO_2 at r.t.p.

8. In $2\,Mg + O_2 \longrightarrow 2\,MgO$, what mass of MgO forms from 4.80 g of Mg (excess O_2)?

9. 25.0 cm^3 of 0.200 $mol\,dm^{-3}$ HCl reacts completely with NaOH. Find the volume of 0.100 $mol\,dm^{-3}$ NaOH required.

Advanced Practice （培优）· 7

Advanced Practice

A. Limiting reagent with leftover gases （限量 + 冷却后余气）

1. Ethene burns: $C_2H_4 + 3\,O_2 \longrightarrow 2\,CO_2 + 2\,H_2O$.
   40.0 cm^3 ethene and 150 cm^3 oxygen are sparked (same T, p), then cooled so water is liquid. Find: (i) limiting reagent (ii) volume of CO_2 (iii) volume of unused O_2.

2. $H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + 2\,H_2O$.
   25.0 cm^3 of 0.200 $mol\,dm^{-3}$ H_2SO_4 is mixed with 30.0 cm^3 of 0.250 $mol\,dm^{-3}$ NaOH. Identify the limiting reagent and calculate n of the salt formed.

B. Yield, purity and formula from data （反推组成）

3. 4.00 g of an impure sample of Mg reacts with excess acid to give 3.36 dm^3 of H_2 at r.t.p.
   $Mg + 2\,HCl \longrightarrow MgCl_2 + H_2$.
   Calculate the percentage purity of the magnesium (assume only Mg produces H_2).

4. A metal M forms chloride MCl_3. 1.35 g of M gives 6.67 g of MCl_3 with excess chlorine. Deduce A_r of M and identify M from the data list.

C. Multi-step solution stoichiometry （气体/滴定反推 c 或式）

5. $2\,H_2O_2 \longrightarrow 2\,H_2O + O_2$.
   25.0 cm^3 of $H_2O_2(aq)$ liberates 48.0 cm^3 O_2 at r.t.p. Find $[H_2O_2]$ in $mol\,dm^{-3}$, then the mass concentration in $g\,dm^{-3}$.

6. 20.0 cm^3 of $Na_2CO_3(aq)$ requires 25.0 cm^3 of 0.200 $mol\,dm^{-3}$ HCl for complete conversion to CO_2 ($Na_2CO_3 + 2\,HCl \longrightarrow 2\,NaCl + H_2O + CO_2$). Find the concentration of the carbonate solution.

7. Ammonia is oxidised: $4\,NH_3 + 5\,O_2 \longrightarrow 4\,NO + 6\,H_2O$.
   A mixture of 120 cm^3 NH_3 and 130 cm^3 O_2 is sparked (same T, p). The reaction does not go to completion. After cooling, the dry gas contains 88.0 cm^3 of NO.
   Determine: (i) the limiting reagent from initial amounts; (ii) the volume of O_2 that actually reacted; (iii) the volumes of unreacted NH_3 and O_2.

3 Chemical Bonding（化学键与结构）

Ionic, covalent and metallic bonding; simple molecular shapes; polarity and electronegativity at AS foundation level.

Rules 规则

- **Ionic bond**（离子键）: electrostatic attraction between oppositely charged ions (often metal + non-metal).
- **Covalent bond**（共价键）: shared pair(s) of electrons between atoms (often non-metal + non-metal).
- **Metallic bond**（金属键）: attraction between positive metal ions and a sea of delocalised electrons.
- **Shapes (VSEPR simple)**（分子构型）: BF_3 trigonal planar; CO_2 linear; CH_4 tetrahedral; NH_3 pyramidal; H_2O bent.
- **Bond angles (approx.)**（键角）: BF_3 120°; CO_2 180°; CH_4 109.5°; NH_3 107°; H_2O 104.5°.
- **Electronegativity**（电负性）: ability of an atom to attract bonding electrons; increases across a period, decreases down a group.
- **Polar covalent bond**（极性键）: bonding electrons shared unequally when electronegativity differs.
- **Polar molecule**（极性分子）: has polar bonds and an asymmetric shape so dipoles do not cancel.

Worked Examples 例题

Example 1. State the type of bonding in NaCl, Cl_2 and Cu.

NaCl: ionic; Cl_2: covalent; Cu: metallic.

Example 2. State the shape and approximate bond angle of CH_4 and of H_2O.

CH_4: tetrahedral, 109.5°.

H_2O: bent (V-shaped), $\approx 104.5°$.

Example 3. Explain briefly why CO_2 is non-polar but H_2O is polar.

Both have polar bonds, but CO_2 is linear so bond dipoles cancel; H_2O is bent so dipoles do not cancel.

Basic Drill （入门综合）· 32

1. Define an ionic bond.

2. Define a covalent bond.

3. Define a metallic bond.

4. State the bonding type in NaCl.

5. State the bonding type in MgO.

6. State the bonding type in H_2.

7. State the bonding type in O_2.

8. State the bonding type in H_2O.

9. State the bonding type in Cu.

10. State the bonding type in Fe.

11. How many bonding pairs surround B in BF_3?

12. State the shape of BF_3.

13. State the approximate bond angle in BF_3.

14. State the shape of CO_2.

15. State the bond angle in CO_2.

16. State the shape of CH_4.

17. State the approximate bond angle in CH_4.

18. State the shape of NH_3.

19. State the approximate bond angle in NH_3.

20. State the shape of H_2O.

21. State the approximate bond angle in H_2O.

22. Why is the bond angle in NH_3 less than in CH_4? (One reason.)

23. Why is the bond angle in H_2O less than in NH_3? (One reason.)

24. Define electronegativity.

25. Which is more electronegative: F or Cl?

26. Which is more electronegative: O or N?

27. Which is more electronegative: Na or Cl?

28. Is the H−Cl bond polar or non-polar?

29. Is the Cl−Cl bond polar or non-polar?

30. Is CO_2 a polar or non-polar molecule?

31. Is H_2O a polar or non-polar molecule?

32. Is CH_4 a polar or non-polar molecule?

Practice （综合练习）· 14

Practice

1. Predict the bonding type in KBr, N_2 and Al.

2. Draw (or describe) the electron arrangement for ionic bonding in NaCl (dot-and-cross idea in words).

3. State the number of lone pairs on N in NH_3 and on O in H_2O.

4. Complete: BF_3 is ____ planar; CO_2 is ____.

5. Rank bond angles: CH_4, NH_3, H_2O from largest to smallest.

6. Explain why BF_3 is non-polar despite polar B–F bonds.

7. Explain why NH_3 is polar.

8. Across Period 3, does electronegativity generally increase or decrease from Na to Cl?

9. Down Group 17, does electronegativity generally increase or decrease?

10. Which bond is more polar: H–F or H–I? Explain briefly.

11. State whether each molecule is polar: BF_3, NH_3, CO_2.

12. Metallic bonding explains why metals conduct electricity. State why, in one sentence.

13. Ionic compounds typically have high melting points. State why, briefly.

14. Simple molecular covalent substances often have low melting points. State why, briefly.

Further Practice （进阶）· 9

Further Practice

A. Bonding types （离子/共价/金属识别）

1. State the type of bonding in NaCl, Cl_2 and Cu, and give a one-line definition of each type.

2. Explain briefly, in terms of electron transfer, how Na and Cl form ionic NaCl.

3. A solid conducts electricity when molten but not when solid, and has a high melting point. Suggest the bonding type.

B. Molecular shapes and bond angles （VSEPR 命名与键角）

4. Predict the shapes and approximate bond angles of $BeCl_2$ and PCl_3.

5. State the shapes of BF_3 and CO_2, and give their bond angles.

6. For CH_4, NH_3 and H_2O, state the shape name and approximate H–X–H angle of each.

C. Electronegativity and simple polarity （极性键与分子极性判断）

7. Using electronegativity ideas, explain why NaCl is ionic but HCl is covalent (polar).

8. State whether CCl_4 is a polar molecule and explain using its shape.

9. State whether NH_3 is polar and explain briefly.

Advanced Practice （培优）· 8

Advanced Practice

A. Shape, angle and electron-pair repulsion （比较 + 必须点名排斥类型）

1. Rank CH_4, NH_3, H_2O by decreasing bond angle and explain using the relative strengths of lp–lp, lp–bp and bp–bp repulsions.

2. Both NH_3 and H_2O are based on four electron pairs around the central atom. Explain why the bond angle in H_2O is smaller than in NH_3.

B. Structure and physical properties （巨共价 / 层状 / 简单分子对比）

3. Compare melting point and electrical conductivity of diamond and graphite. Explain both differences using bonding and structure.

4. Explain why SiO_2 has a much higher melting point than CO_2, even though both contain covalent bonds between non-metals.

5. Iodine sublimes on gentle heating. Explain this in terms of the particles and forces that must be overcome (contrast with melting NaCl).

C. Deducing bonding and polarity （物性表 → 键型；极性易错）

6. A white solid has high m.p., conducts only when molten, and gives a conducting aqueous solution. Another solid has very high m.p. and never conducts. Identify the bonding/structure type of each.

7. BF_3 has polar bonds but is a non-polar molecule, while NF_3 is polar. Explain the difference using shape and bond dipoles.

8. Four samples are tested for electrical conductivity: solid Al, molten Al, solid Al_2O_3, molten Al_2O_3. For each sample, state whether it conducts and identify the mobile charge carrier (if any). Explain why "contains ions" does not always mean "conducts".

4 States of Matter（物态）

Ideal gas law basics; intermolecular forces; types of solid structure at AS foundation level.

Rules 规则

- **Ideal gas equation**（理想气体方程）：$pV = nRT$; use consistent units (p in Pa, V in m^3, T in K, $R = 8.31\,\mathrm{J\,K^{-1}\,mol^{-1}}$).
- **r.t.p. shortcut**（室温常压近似）：$V_\mathrm{m} = 24.0\,\mathrm{dm^3\,mol^{-1}}$ at r.t.p. (when stated).
- **Kelvin**（热力学温度）：$T/\mathrm{K} = t/^\circ\mathrm{C} + 273$.
- **van der Waals / id–id**（瞬时偶极–诱导偶极）：present in all molecules; stronger for larger / more polarisable species.
- **Permanent dipole–dipole**（永久偶极–偶极）：between polar molecules.
- **Hydrogen bonding**（氢键）：H bonded to N, O or F interacting with lone pair on N/O/F.
- **Solid types**（固体类型）：ionic, metallic, simple molecular, giant covalent (e.g. diamond, graphite, $\mathrm{SiO_2}$).
- **Melting / boiling**（熔沸点）：depend on strength of forces between particles (not covalent bonds inside simple molecules).

Worked Examples 例题

Example 1. What volume does 0.500 mol of an ideal gas occupy at r.t.p.?

$V = 0.500 \times 24.0 = 12.0\,\mathrm{dm^3}$.

Example 2. Convert 25°C to kelvin. State the strongest intermolecular force in liquid water.

$T = 25 + 273 = 298\,\mathrm{K}$.

Hydrogen bonding.

Example 3. Classify the solid type of NaCl, diamond and ice ($\mathrm{H_2O}$).

NaCl: ionic; diamond: giant covalent; ice: simple molecular (with H-bonding).

Basic Drill （入门综合）· 28

1. Write the ideal gas equation.

2. Convert 0°C to kelvin.

3. Convert 100°C to kelvin.

4. What volume at r.t.p. is occupied by 1.00 mol of gas?

5. What volume at r.t.p. is occupied by 0.250 mol of O_2?

6. What amount of gas occupies 12.0 dm^3 at r.t.p.?

7. What amount of gas occupies 6.00 dm^3 at r.t.p.?

8. Find n in 48.0 dm^3 of N_2 at r.t.p.

9. Find the volume at r.t.p. of 0.100 mol CO_2.

10. State Avogadro's law in one sentence (same T, p).

11. Name the intermolecular force present in all molecules.

12. Name the strongest intermolecular force in $HF(l)$.

13. Name the strongest intermolecular force in $HCl(l)$.

14. Name the strongest intermolecular force in $CH_4(l)$.

15. Which has hydrogen bonding: H_2O or H_2S?

16. Which has hydrogen bonding: NH_3 or PH_3?

17. Classify solid NaCl: ionic / metallic / simple molecular / giant covalent.

18. Classify solid Cu: ionic / metallic / simple molecular / giant covalent.

19. Classify solid I_2: ionic / metallic / simple molecular / giant covalent.

20. Classify diamond: ionic / metallic / simple molecular / giant covalent.

21. Classify graphite: ionic / metallic / simple molecular / giant covalent.

22. Classify solid SiO_2: ionic / metallic / simple molecular / giant covalent.

23. Why do metals conduct electricity in the solid state? (One reason.)

24. Why do ionic solids not conduct when solid but do when molten?

25. Why does ice melt at a relatively high temperature for its M_r?

26. Which has the higher boiling point: H_2O or H_2S?

27. Which solid has a higher melting point: NaCl or I_2?

28. At constant T, if p doubles and n is fixed, what happens to V (ideal gas)?

Practice （综合练习）· 12

Practice

1. Calculate the volume at r.t.p. of 4.40 g of CO_2.

2. A gas sample occupies $2.40\,\text{dm}^3$ at r.t.p. Find n and the number of molecules ($L = 6.02 \times 10^{23}\,\text{mol}^{-1}$).

3. Using $pV = nRT$, find n if $p = 100\,\text{kPa}$, $V = 2.00\,\text{dm}^3$, $T = 300\,\text{K}$.
   (Hint: $100\,\text{kPa} = 1.00 \times 10^5\,\text{Pa}$; $2.00\,\text{dm}^3 = 2.00 \times 10^{-3}\,\text{m}^3$; $R = 8.31\,\text{J K}^{-1}\,\text{mol}^{-1}$.)

4. Rank intermolecular force strength typically: H-bond, id–id (small molecule), permanent dipole–dipole.

5. Explain why Br_2 has a higher boiling point than Cl_2.

6. Explain why NH_3 has a higher boiling point than PH_3.

7. Compare melting points qualitatively: diamond and graphite versus I_2. Give a brief reason for the much lower melting point of I_2.

8. State two differences between ionic and metallic solids (structure / properties).

9. 0.200 mol of gas is heated from 300 K to 600 K at constant p. What happens to volume?

10. A mixture of 0.100 mol H_2 and 0.100 mol O_2 is at r.t.p. Find the total volume (assume ideal, no reaction).

11. Explain why graphite conducts electricity but diamond does not.

12. State whether liquid CCl_4 has H-bonding, and name its main intermolecular force.

Further Practice （进阶）· 8

Further Practice

A. Ideal gas and r.t.p. volumes （$V = n \times 24.0$ 基础）

1. 3.20 g of O_2 occupies what volume at r.t.p.? Also find n.

2. Calculate the volume at r.t.p. of 0.250 mol of N_2.

3. Equal volumes of N_2 and H_2 at the same T, p contain equal numbers of molecules. Name

this law / idea.

B. Types of solid （结构类型识别）

4. State the structure type of diamond, ice, and sodium chloride (giant covalent / simple molecular / ionic).

5. Which type of solid typically has a low melting point and does not conduct: ionic, giant covalent, or simple molecular?

6. Give one example of a metallic solid and state what carries the current when it conducts.

C. Intermolecular forces (naming) （作用力类型）

7. Name the strongest intermolecular force in H_2O, in HCl, and in Br_2.

8. State whether boiling point generally increases or decreases as the strength of intermolecular forces increases.

Advanced Practice （培优）· 8

Advanced Practice

A. $pV = nRT$ with stoichiometry （先算 n，再接反应）

1. A sample of CO_2 has $p = 100\,\text{kPa}$, $V = 2.49\,\text{dm}^3$, $T = 300\,\text{K}$ ($R = 8.31$). Calculate n, then the volume this sample would occupy at r.t.p.

2. 0.240 g of Mg reacts with excess acid in a gas syringe at 101 kPa and 298 K.
$\text{Mg} + 2\,\text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$.
Calculate the expected volume of H_2 using $pV = nRT$ ($R = 8.31$).

B. Boiling points and IMF （同电子数对比 / 氢键异常）

3. C_2H_6 and CH_3OH have similar M_r. Explain why methanol has the much higher boiling point.

4. Explain why boiling points of noble gases increase down the group, and why HF has a higher boiling point than HCl despite HCl having more electrons.

C. Structure and kinetic theory ideas （结构推断 + 分子运动论）

5. Solid A: high m.p., conducts when molten only. Solid B: very high m.p., never conducts. Solid C: low m.p., never conducts. Assign ionic / giant covalent / simple molecular to A–C and justify with forces overcome on melting.

6. Two equal flasks hold CH_4 and SO_2 at the same T and p. Compare numbers of molecules and densities. Explain.

7. Explain why ice floats on water using hydrogen bonding and open lattice structure, and state what this implies about the average distance between molecules in ice vs liquid water.

8. A sealed vessel contains a real gas. Step 1: the volume is halved at constant T. Step 2: the gas is then heated while the volume stays at that halved value.
   (i) Using kinetic theory, explain the effect of each step on the pressure.
   (ii) State two assumptions of the ideal-gas model that become less valid after Step 1, and explain why each assumption fails.

5 Chemical Energetics（热化学）

Enthalpy change definitions; simple Hess cycles; bond energies; $q = mc\Delta T$ calorimetry basics.

Rules 规则

- **Enthalpy change** ΔH（焓变）：heat change at constant pressure; exothermic $\Delta H < 0$; endothermic $\Delta H > 0$.
- **Standard enthalpy of formation** ΔH_f°（生成焓）: enthalpy change when 1 mol of compound forms from its elements in standard states.
  e.g. $\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)} \quad \Delta H = \Delta H_f^\circ(\mathrm{CO_2})$.
- **Standard enthalpy of combustion** ΔH_c°（燃烧焓）：enthalpy change when 1 mol of substance burns completely in oxygen under standard conditions.
  e.g. $\mathrm{CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(l)} \quad \Delta H = \Delta H_c^\circ(\mathrm{CH_4})$.
- **Hess's law**（盖斯定律）：overall ΔH is independent of route; sum of ΔH around a cycle is zero.
- **Bond energy**（键能）：average energy to break 1 mol of a given bond in gaseous molecules; bond breaking endothermic.
  e.g. $\mathrm{Cl_2(g) \longrightarrow 2\,Cl(g)}$; $\mathrm{H_2(g) \longrightarrow 2\,H(g)}$.
- **Reaction from bond energies**（由键能估 ΔH）：$\Delta H \approx \sum$(bonds broken) $-$ $\sum$(bonds formed).
- **Calorimetry**（量热）：$q = mc\Delta T$; for solutions often take $c = 4.18\,\mathrm{J\,g^{-1}\,^\circ C^{-1}}$ and density $\approx 1\,\mathrm{g\,cm^{-3}}$.
- ΔH **from** q（由热量求焓变）: $\Delta H = -\dfrac{q}{n}$ for exothermic processes if heat released by reaction $= q$ absorbed by surroundings (sign carefully).
- **Energy profile**（能量剖面）：plot enthalpy vs reaction pathway; E_a is the energy barrier; $\Delta H = H_{\text{products}} - H_{\text{reactants}}$.

Worked Examples 例题

Example 1. Is combustion of methane typically exothermic or endothermic? What is the sign of ΔH?

Exothermic; ΔH is negative.

Example 2. 50.0 g of water is heated by 4.00°C. Calculate q using $c = 4.18\,\mathrm{J\,g^{-1}\,^\circ C^{-1}}$.

$q = mc\Delta T = 50.0 \times 4.18 \times 4.00 = 836\,\mathrm{J}$.

Example 3. Using bond energies: break 1 H – H and 1 Cl – Cl; form 2 H – Cl.
Data: H – H 436, Cl – Cl 242, H – Cl 431 (all $\mathrm{kJ\,mol^{-1}}$). Estimate ΔH.

Broken: $436 + 242 = 678$.

Formed: $2 \times 431 = 862$.

$\Delta H = 678 - 862 = -184\,\mathrm{kJ\,mol^{-1}}$.

Basic Drill （入门综合）· 32

1. Define enthalpy change, ΔH.

2. State the sign of ΔH for an exothermic reaction.

3. State the sign of ΔH for an endothermic reaction.

4. Define standard enthalpy of formation, $\Delta H_{\mathrm{f}}^{\circ}$.

5. Define standard enthalpy of combustion, $\Delta H_{\mathrm{c}}^{\circ}$.

6. What is $\Delta H_{\mathrm{f}}^{\circ}$ of an element in its standard state?

7. Is bond breaking endothermic or exothermic?

8. Is bond forming endothermic or exothermic?

9. Write the word equation idea of Hess's law in one sentence.

10. Write the formula $q = mc\Delta T$ and name each symbol.

11. Calculate q when 100 g water rises by 2.00°C ($c = 4.18\,\mathrm{J\,g^{-1}\,°C^{-1}}$).

12. Calculate q when 25.0 g water rises by 8.00°C ($c = 4.18$).

13. Convert 836 J to kJ.

14. Convert 2.50 kJ to J.

15. A reaction releases 500 J to warm water. Is the reaction exo- or endothermic?

16. If q (to water) = 1.00 kJ and n(fuel) = 0.0100 mol, estimate ΔH in $\mathrm{kJ\,mol^{-1}}$ (exothermic combustion).

17. State whether ΔH_c° of carbon is typically positive or negative.

18. In a Hess cycle, route 1 has $\Delta H = +10$ then -40. What is overall ΔH?

19. Bonds broken cost $800\,\mathrm{kJ\,mol^{-1}}$; bonds formed release $1000\,\mathrm{kJ\,mol^{-1}}$. Estimate ΔH.

20. Bonds broken 600; bonds formed 450 (same units $\mathrm{kJ\,mol^{-1}}$). Estimate ΔH.

21. For $\mathrm{H_2(g) + Cl_2(g) \longrightarrow 2\,HCl\,(g)}$, list which bonds break and which form.

22. Why are bond energies only approximate for ΔH of reactions? (One reason.)

23. Standard conditions for enthalpy data: typical T and p values used at AS.

24. Combustion of $0.0200\,\mathrm{mol}$ methanol releases $14.0\,\mathrm{kJ}$ to the surroundings. Find ΔH_c (approx).

25. $m = 50.0\,\mathrm{g}$, $c = 4.18$, $\Delta T = 10.0^\circ\mathrm{C}$. Find q in kJ.

26. Using Q25, if this heat came from burning $0.0100\,\mathrm{mol}$ of fuel, find ΔH in $\mathrm{kJ\,mol^{-1}}$.

27. State two common heat-loss errors in simple calorimetry.

28. Endothermic reaction in a beaker: does the beaker feel hotter or colder?

29. Exothermic dissolving: sign of ΔH_sol?

30. Formation of $\mathrm{CO_2(g)}$ from C(graphite) and $\mathrm{O_2}$: related to which standard enthalpy?

31. If $\Delta H_1 = +50$ and $\Delta H_2 = -20$ are two steps of a Hess route, overall $\Delta H =$?

32. True or false: stronger bonds forming than breaking $\Rightarrow$ usually exothermic overall.

Practice （综合练习）· 14

Practice

1. $75.0\,\mathrm{cm^3}$ of water ($75.0\,\mathrm{g}$) warms by $6.40^\circ\mathrm{C}$. Calculate q ($c = 4.18\,\mathrm{J\,g^{-1}\,^\circ C^{-1}}$) in kJ.

2. Neutralisation: $50.0\,\mathrm{cm^3}$ of $1.00\,\mathrm{mol\,dm^{-3}}$ HCl mixes with $50.0\,\mathrm{cm^3}$ of $1.00\,\mathrm{mol\,dm^{-3}}$ NaOH; $\Delta T = +6.50^\circ\mathrm{C}$.
   Assume total mass $100\,\mathrm{g}$, $c = 4.18$. Find q and ΔH per mole of water formed.

3. Bond energies / $\mathrm{kJ\,mol^{-1}}$: C–H 410, Cl–Cl 242, C–Cl 340, H–Cl 431.
   For $\mathrm{CH_4 + Cl_2 \longrightarrow CH_3Cl + HCl}$, estimate ΔH (break one C–H and Cl–Cl; form C–Cl and H–Cl).

4. Given $\Delta H_\mathrm{f}^\circ(\mathrm{CO_2}) = -394$, $\Delta H_\mathrm{f}^\circ(\mathrm{H_2O}) = -286$, $\Delta H_\mathrm{f}^\circ(\mathrm{CH_4}) = -75$ ($\mathrm{kJ\,mol^{-1}}$),
   find $\Delta H_\mathrm{c}^\circ$ for $\mathrm{CH_4}$: $\mathrm{CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O}$.

5. Hess: ΔH for A→B is unknown. A→C = +30; C→B = −80 ($\mathrm{kJ\,mol^{-1}}$). Find A→B.

6. Combustion of $0.500\,\mathrm{g}$ ethanol ($M_\mathrm{r} = 46.0$) raises $200\,\mathrm{g}$ water by $8.00^\circ\mathrm{C}$ ($c = 4.18$). Estimate ΔH_c in $\mathrm{kJ\,mol^{-1}}$.

7. State why experimental $|\Delta H_\mathrm{c}|$ is often smaller than data-book values.

8. $\Delta H_\mathrm{f}^\circ(\mathrm{H_2O\,(l)}) = -286\,\mathrm{kJ\,mol^{-1}}$. Write the formation equation.

9. Using bond energies: N≡N 945, H–H 436, N–H 390.
   Estimate ΔH for $\mathrm{N_2 + 3\,H_2 \longrightarrow 2\,NH_3}$.

10. A reaction has $\Delta H = +45\,\text{kJ mol}^{-1}$. Is heat absorbed or released by the system?

11. State Hess's law in symbols for a closed cycle of three enthalpy changes summing to zero.

12. $q = 2.09\,\text{kJ}$ absorbed by water from a reaction of $0.0500\,\text{mol}$ reactant (1:1). Find ΔH.

13. Rank typically most exothermic process: bond forming in products much stronger vs much weaker than bonds broken.

14. Calculate ΔT if $q = 1672\,\text{J}$, $m = 100\,\text{g}$, $c = 4.18\,\text{J g}^{-1}\,^\circ\text{C}^{-1}$.

Further Practice （进阶）· 9

Further Practice

A. Enthalpy change basics （符号 · 定义）

1. Define standard enthalpy change of combustion, ΔH_c°.

2. State whether combustion of methane is exothermic or endothermic, and give the sign of ΔH.

3. A Hess cycle forms a closed triangle with $\Delta H_1 = +120$, $\Delta H_2 = -80$, and $\Delta H_1 + \Delta H_2 + \Delta H_3 = 0$. Find ΔH_3.

B. Bond energy estimates （断键吸热 · 成键放热）

4. Bond energies / kJ mol^{-1}: H–H 436, Cl–Cl 242, H–Cl 431. Estimate ΔH for $\text{H}_2 + \text{Cl}_2 \longrightarrow 2\,\text{HCl}$.

5. Bond energies: C=C 610, H–H 436, C–C 350, C–H 410.
   Estimate ΔH for $C_2H_4 + H_2 \longrightarrow C_2H_6$ (ethene: one C=C and 4 C–H; ethane: one C–C and 6 C–H; also break H–H).

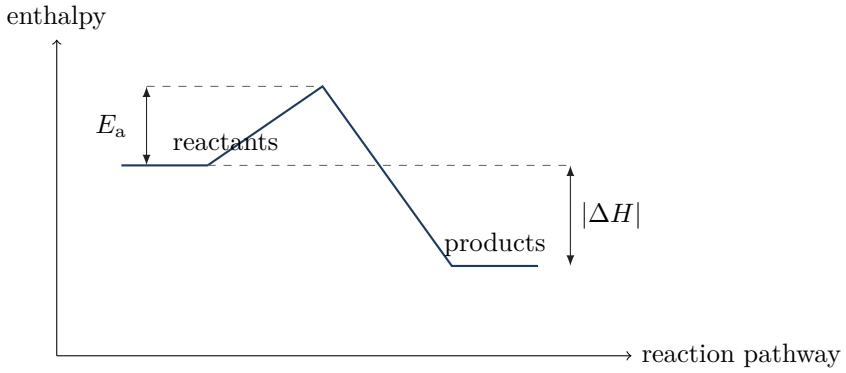
C. Simple calorimetry （$q = mc\Delta T$ 单步）

6. 50.0 g of water is heated by 8.00°C ($c = 4.18\,\mathrm{J\,g^{-1}\,°C^{-1}}$). Calculate q in J and kJ.

7. 0.0200 mol of a fuel releases enough heat to raise 100 g water by 16.0°C ($c = 4.18$). Calculate q and ΔH of combustion (assume all heat to water).

8. State the equation $q = mc\Delta T$ and identify each symbol.

D. Energy profile diagram （能量剖面图）



9. For the energy profile shown:
   (i) State whether the reaction is exothermic or endothermic, and the sign of ΔH.
   (ii) Mark (or describe) what E_a and $|\Delta H|$ represent on the diagram.
   (iii) Sketch how the profile changes if a catalyst is added (qualitative).

Advanced Practice （培优）· 9

Advanced Practice

A. Calorimetry to ΔH (non-obvious n) （先定反应量再求 ΔH）

1. $25.0\,\mathrm{cm^3}$ of $0.400\,\mathrm{mol\,dm^{-3}}$ $CuSO_4$ reacts with excess Zn.
   Solution mass $= 25.0\,\mathrm{g}$, $c = 4.18$, $\Delta T = +6.00^\circ\mathrm{C}$.
   Calculate ΔH for $\mathrm{Zn} + \mathrm{Cu^{2+}} \longrightarrow \mathrm{Zn^{2+}} + \mathrm{Cu}$, and state why using excess Zn matters.

2. $40.0\,\mathrm{cm^3}$ of $0.500\,\mathrm{mol\,dm^{-3}}$ HCl is mixed with $60.0\,\mathrm{cm^3}$ of $0.500\,\mathrm{mol\,dm^{-3}}$ NaOH in a cup; $\Delta T = +4.20^\circ\mathrm{C}$. Take total mass $= 100\,\mathrm{g}$, $c = 4.18$. Identify the limiting reagent for neutralisation and find ΔH per mole of water formed.

B. Hess cycles requiring route design （燃烧焓 ↔ 生成焓）

3. Given $\Delta H_c^\circ(\mathrm{C}) = -394$, $\Delta H_c^\circ(\mathrm{H_2}) = -286$, $\Delta H_c^\circ(\mathrm{CH_3OH}) = -726$ $(\mathrm{kJ\,mol^{-1}})$, calculate $\Delta H_f^\circ(\mathrm{CH_3OH})$. State the formation equation for methanol that this value refers to.

4. ΔH_f°: $\mathrm{C_2H_2}$=+227, $\mathrm{CO_2}$=-394, $\mathrm{H_2O\,(l)}$=-286 $(\mathrm{kJ\,mol^{-1}})$. Calculate ΔH_c° for the complete combustion of 1 mol of ethyne to $\mathrm{CO_2}$ and $\mathrm{H_2O\,(l)}$.

C. Bond energies and experimental interpretation （平均值近似 + 误差方向）

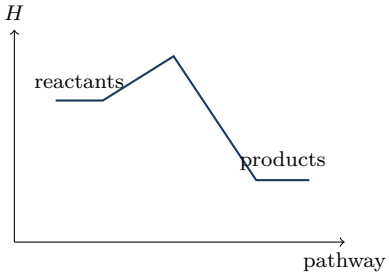
5. Using bond energies $\mathrm{N_2}$ 944 (triple bond), H – H 436, N – H 388 $(\mathrm{kJ\,mol^{-1}})$, estimate ΔH for $\mathrm{N_2 + 3\,H_2 \longrightarrow 2\,NH_3}$, and explain one reason the experimental ΔH differs.

6. In a spirit-burner calorimetry experiment, consider three separate errors: (a) heat is lost to the surroundings; (b) combustion is incomplete (some CO forms); (c) water condenses on the burner before the final mass is recorded.
   For each, state the effect on the calculated $|\Delta H_c|$ and explain briefly.

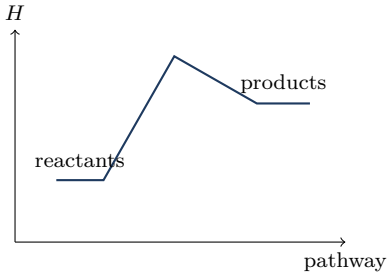
7. Bond-energy calculations assume gaseous molecules and average bond enthalpies. Explain why this makes ΔH for $H_2O(l)$ formation from bond energies unreliable without extra data.

8. An endothermic dissolution gives a temperature fall. Starting from $q = mc\Delta T$, explain how to obtain a positive ΔH per mole of solute, including the sign of q.

D. Energy profiles (exo vs endo) （放热/吸热剖面对比）



A (exothermic)



B (endothermic)

9. Profiles A and B are shown.
   (i) On each sketch, identify which quantity is larger: E_a(forward) or E_a(reverse).
   (ii) For profile B, state the sign of ΔH and explain using reactant/product enthalpy levels.
   (iii) A catalyst is added to A. State the effect on ΔH and on E_a(forward).

6 Electrochemistry（电化学）

Oxidation numbers; identifying redox; simple half-equations; electrolysis product prediction at foundation level.

Rules 规则

- **Oxidation number**（氧化数）: bookkeeping charge on an atom in a compound/ion; elements in elemental form = 0.
- **Common ON rules**（常用规则）: Group 1 = +1; Group 2 = +2; O usually −2 (peroxides −1); H usually +1 (hydrides −1); F always −1.
- **Oxidation / reduction**（氧化/还原）: oxidation: loss of electrons / ON increases; reduction: gain of electrons / ON decreases.
- **Oxidising / reducing agent**（氧化剂/还原剂）: oxidising agent is reduced; reducing agent is oxidised.
- **Half-equations**（半反应）: show e^- explicitly; balance atoms and charge.
  e.g. $Zn \longrightarrow Zn^{2+} + 2\,e^-$; $Cu^{2+} + 2\,e^- \longrightarrow Cu$.
- **Electrolysis**（电解）: decomposition by electric current; cathode reduction; anode oxidation.
- **Aqueous electrolysis (simple)**（水溶液电解简化）: at cathode: less reactive metal or H_2 from H^+/H_2O; at anode: halide or O_2 from OH^-/H_2O (concentration matters for dilute vs conc. halide).
  common half-equations:
  cathode: $2\,H^+ + 2\,e^- \longrightarrow H_2$ or $2\,H_2O + 2\,e^- \longrightarrow H_2 + 2\,OH^-$
  anode: $2\,Cl^- \longrightarrow Cl_2 + 2\,e^-$ or $4\,OH^- \longrightarrow O_2 + 2\,H_2O + 4\,e^-$.
- **Molten salts**（熔融盐）: cathode: metal; anode: non-metal.
  e.g. molten NaCl: $Na^+ + e^- \longrightarrow Na$; $2\,Cl^- \longrightarrow Cl_2 + 2\,e^-$.

Worked Examples 例题

Example 1. Find the oxidation number of S in SO_4^{2-} and of Mn in MnO_4^-.

In SO_4^{2-}: $x + 4(-2) = -2 \Rightarrow x = +6$.

In MnO_4^-: $x + 4(-2) = -1 \Rightarrow x = +7$.

Example 2. In $2\,Mg + O_2 \longrightarrow 2\,MgO$, which species is oxidised and which is reduced?

Mg is oxidised (0 to +2); O_2 is reduced (0 to −2).

Example 3. State the products at cathode and anode for electrolysis of molten NaCl.

Cathode: Na; anode: Cl_2.

Basic Drill （入门综合）· 30

1. State the oxidation number of an element in its elemental form.

2. State the usual oxidation number of O in compounds.

3. State the usual oxidation number of H in compounds.

4. Find ON of Na in NaCl.

5. Find ON of Mg in MgO.

6. Find ON of Cl in NaCl.

7. Find ON of C in CO_2.

8. Find ON of C in CH_4.

9. Find ON of N in NH_3.

10. Find ON of N in NO_3^-.

11. Find ON of S in H_2SO_4.

12. Find ON of Cr in $Cr_2O_7{}^{2-}$.

13. Define oxidation in terms of electrons.

14. Define reduction in terms of electrons.

15. Define oxidation in terms of oxidation number.

16. In $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$, which species is oxidised?

17. In the reaction of Q16, which species is reduced?

18. Identify the reducing agent in Q16.

19. Identify the oxidising agent in Q16.

20. Write the half-equation for $Zn \longrightarrow Zn^{2+}$ (include e^-).

21. Write the half-equation for $Cu^{2+} \longrightarrow Cu$.

22. Write the half-equation for $2\,Cl^- \longrightarrow Cl_2$.

23. Write the half-equation for $2\,H^+ \longrightarrow H_2$.

24. In electrolysis, does reduction occur at the cathode or anode?

25. Products of molten $PbBr_2$ electrolysis: cathode and anode.

26. Dilute aqueous H_2SO_4 electrolysis: gas at cathode?

27. Dilute aqueous H_2SO_4 electrolysis: gas at anode?

28. Concentrated aqueous NaCl electrolysis: gas commonly at anode?

29. Aqueous $CuSO_4$ with Cu electrodes: what happens to anode mass (qualitative)?

30. Is $Fe^{2+} \longrightarrow Fe^{3+} + e^-$ oxidation or reduction?

Practice （综合练习）· 14

Practice

1. Find ON of each atom in $K_2Cr_2O_7$.

2. Find ON of Cl in Cl_2, Cl^-, and $ClO_3{}^-$.

3. Balance the ionic half-equation: $Fe^{2+} \longrightarrow Fe^{3+}$ (already almost done — add e^-).

4. Combine: $\mathrm{Mg} \longrightarrow \mathrm{Mg^{2+}} + 2\,\mathrm{e^-}$ and $2\,\mathrm{H^+} + 2\,\mathrm{e^-} \longrightarrow \mathrm{H_2}$ into an overall equation.

5. In $\mathrm{Cl_2} + 2\,\mathrm{Br^-} \longrightarrow 2\,\mathrm{Cl^-} + \mathrm{Br_2}$, identify oxidising and reducing agents.

6. Show ON changes for C in $\mathrm{C} + \mathrm{O_2} \longrightarrow \mathrm{CO_2}$ and classify redox.

7. Electrolysis of molten $\mathrm{Al_2O_3}$ (simplified): products at cathode and anode.

8. Aqueous NaCl (dilute) vs concentrated: which anode product is more likely for dilute?

9. Write a balanced half-equation for $\mathrm{MnO_4^-} \longrightarrow \mathrm{Mn^{2+}}$ in acid (include $\mathrm{H^+}$, $\mathrm{H_2O}$, $\mathrm{e^-}$).

10. Write a balanced half-equation for $\mathrm{Cr_2O_7^{2-}} \longrightarrow \mathrm{Cr^{3+}}$ in acid.

11. Predict cathode product for aqueous $\mathrm{AgNO_3}$ (inert electrodes).

12. Predict anode product for aqueous KI (concentrated, inert electrodes).

13. Explain briefly why solid ionic compounds do not electrolyse but molten ones can.

14. In electrolysis of aqueous $\mathrm{CuSO_4}$ with inert electrodes, state cathode and anode products.

Further Practice （进阶）· 9

Further Practice

A. Oxidation numbers （氧化数赋值）

1. Assign the oxidation number of S in SO_2, SO_4^{2-} and H_2S.

2. Assign ON of N in NH_3, NO and HNO_3.

3. Assign ON of Cr in $Cr_2O_7^{2-}$ and of Mn in MnO_4^-.

B. Redox identification and half-equations （氧化还原识别与半方程）

4. In $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$, identify what is oxidised and what is reduced, and write both half-equations.

5. Balance in acid: $MnO_4^- \longrightarrow Mn^{2+}$ (include H^+, H_2O, e^-).

6. Balance in acid: $Cr_2O_7^{2-} \longrightarrow Cr^{3+}$.

C. Electrolysis products (standard cases) （电解产物判断）

7. State cathode and anode products for molten NaCl (inert electrodes).

8. State cathode and anode products for dilute aqueous H_2SO_4 (inert electrodes).

9. For concentrated aqueous NaCl, state the usual anode product and explain briefly why it differs from dilute solution.

Advanced Practice （培优）· 8

Advanced Practice

A. Average ON and disproportionation （平均氧化数与歧化）

1. Assign the *average* ON of S in $S_2O_3^{2-}$ and in $S_4O_6^{2-}$. Explain what "average" means here.

2. For $2\,H_2O_2 \longrightarrow 2\,H_2O + O_2$, show using oxidation numbers that oxygen is both oxidised and reduced.

B. Combining half-equations and electrolysis stoichiometry （半方程合并与电量计量）

3. Construct the overall ionic equation from
$Fe^{2+} \longrightarrow Fe^{3+} + e^-$ and $MnO_4^- + 8\,H^+ + 5\,e^- \longrightarrow Mn^{2+} + 4\,H_2O$.

4. In electrolysis of molten NaCl, 0.200 mol of electrons passes. Calculate maximum $n(\mathrm{Na})$ and $n(\mathrm{Cl_2})$.

C. Aqueous electrolysis reasoning （水溶液优先放电解释）

5. Explain why electrolysis of aqueous Na_2SO_4 gives H_2 and O_2, not sodium metal or sulfur.

6. Dilute H_2SO_4 with inert electrodes: write cathode and anode half-equations, and state the overall cell reaction.

7. Aqueous $CuSO_4$ with *copper* electrodes: state what happens at each electrode and explain why the blue colour intensity is approximately unchanged.

8. The same charge (0.200 mol of electrons) is passed through (i) molten NaCl and (ii) concentrated aqueous NaCl, both with inert electrodes. State cathode and anode products in each case, and calculate the amount of each electrode product formed.

7 Equilibria（化学平衡）

Dynamic equilibrium; Le Chatelier's principle; writing K_c and simple calculations; AS acid–base basics.

Rules 规则

- **Dynamic equilibrium**（动态平衡）: forward and reverse rates equal; concentrations constant in a closed system.
- **Le Chatelier**（勒夏特列）: system opposes a change in concentration, pressure (gas) or temperature.
- **Temperature**（温度）: increase T favours endothermic direction; decrease T favours exothermic direction.
- **Pressure (gases)**（压强）: increase p favours side with fewer gas moles.
- **Concentration**（浓度）: adding a species shifts equilibrium to consume it.
- **Equilibrium constant** K_c（平衡常数）: for $aA + bB \rightleftharpoons cC + dD$: $K_c = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ (homogeneous).

  e.g. Haber: $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$, $K_c = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$.
- **Catalyst**（催化剂）: speeds both directions equally; no change to equilibrium position / K_c.
- **Acid–base (AS)**（酸碱）: Brønsted–Lowry: acid donates H^+, base accepts H^+;
  strong: $HCl \longrightarrow H^+ + Cl^-$ (or $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$);
  weak: $HA \rightleftharpoons H^+ + A^-$ (e.g. $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$);
  water: $H_2O \rightleftharpoons H^+ + OH^-$, $K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$ at 25°C.

Worked Examples 例题

Example 1. For $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$ ($\Delta H < 0$), state the effect of increasing pressure on yield of NH_3.

Equilibrium shifts right (fewer gas moles); yield of NH_3 increases.

Example 2. Write K_c for $H_2 + I_2 \rightleftharpoons 2\,HI$.

$K_c = \dfrac{[HI]^2}{[H_2][I_2]}$.

Example 3. At equilibrium $[H_2] = 0.20$, $[I_2] = 0.20$, $[HI] = 1.60\,\text{mol dm}^{-3}$. Calculate K_c for Q2's reaction.

$$K_c = \frac{(1.60)^2}{(0.20)(0.20)} = 64.$$

Basic Drill （入门综合）· 32

1. Define dynamic equilibrium.

2. State two features of a system at equilibrium.

3. Write K_c for $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$.

4. Write K_c for $H_2 + I_2 \rightleftharpoons 2\,HI$.

5. Write K_c for $PCl_5 \rightleftharpoons PCl_3 + Cl_2$.

6. Does a catalyst change K_c?

7. Does a catalyst change the position of equilibrium?

8. For an exothermic forward reaction, increasing T shifts equilibrium which way?

9. For an endothermic forward reaction, increasing T shifts equilibrium which way?

10. $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$: effect of increasing pressure on equilibrium position?

11. $H_2 + I_2 \rightleftharpoons 2\,HI$: effect of increasing pressure?

12. $PCl_5 \rightleftharpoons PCl_3 + Cl_2$: effect of increasing pressure?

13. Adding extra H_2 to Haber equilibrium shifts which way?

14. Removing NH_3 from Haber equilibrium shifts which way?

15. State Le Chatelier's principle in one sentence.

16. Define a Brønsted–Lowry acid.

17. Define a Brønsted–Lowry base.

18. Is HCl a strong or weak acid in water?

19. Is CH_3COOH a strong or weak acid in water?

20. Write the dissociation equation for HCl in water.

21. Write the dissociation equilibrium for ethanoic acid in water.

22. What is $[H^+]$ in $0.0100\,mol\,dm^{-3}$ HCl (strong, fully dissociated)?

23. What is $[OH^-]$ in $0.0100\,\mathrm{mol\,dm^{-3}}$ NaOH (strong)?

24. At 25°C, $K_w = 1.00 \times 10^{-14}$. Find $[OH^-]$ if $[H^+] = 1.00 \times 10^{-3}\,\mathrm{mol\,dm^{-3}}$.

25. Find $[H^+]$ if $[OH^-] = 1.00 \times 10^{-2}\,\mathrm{mol\,dm^{-3}}$ at 25°C.

26. Neutral water at 25°C: $[H^+] =$?

27. True or false: at equilibrium, reactants have stopped changing into products.

28. True or false: K_c is constant at constant temperature.

29. If K_c is very large, are products or reactants favoured?

30. If K_c is very small, are products or reactants favoured?

31. For $2\,SO_2 + O_2 \rightleftharpoons 2\,SO_3$ (exothermic), how to increase SO_3 yield (two conditions)?

32. State the effect of increasing concentration of a solid on K_c (heterogeneous — solids omitted): does K_c change?

Practice （综合练习）· 14

Practice

1. For $H_2 + I_2 \rightleftharpoons 2\,HI$, equilibrium amounts in $1.00\,dm^3$: $n(H_2) = 0.40$, $n(I_2) = 0.40$, $n(HI) = 2.40$. Calculate K_c.

2. Same reaction: if $[H_2] = [I_2] = 0.10$ and $K_c = 64$, find $[HI]$.

3. $N_2O_4 \rightleftharpoons 2\,NO_2$ (endothermic forward). State effects of higher T and higher p on the equilibrium amount of NO_2.

4. Haber: why is a compromise temperature used industrially? (One yield reason + one rate reason.)

5. Write K_c for $2\,SO_2 + O_2 \rightleftharpoons 2\,SO_3$.

6. Equilibrium: $[SO_2] = 0.20$, $[O_2] = 0.10$, $[SO_3] = 0.40$ (all $mol\,dm^{-3}$). Find K_c for Q5.

7. Adding a catalyst to Haber process: effect on rate and on equilibrium yield.

8. Identify acid and base in $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$ (Brønsted–Lowry).

9. Identify conjugate acid–base pairs in Q8.

10. Calculate $[H^+]$ and $[OH^-]$ in $0.100\,mol\,dm^{-3}$ HNO_3 at 25°C.

11. A weak acid HA is only partly dissociated. Compare $[H^+]$ with the formal acid concentration (qualitative).

12. For $PCl_5 \rightleftharpoons PCl_3 + Cl_2$, starting with only PCl_5, describe how amounts change until equilibrium.

13. State whether decreasing volume (increasing p) favours reactants or products for $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$.

14. $K_c = 4.0$ for $A \rightleftharpoons 2\,B$ in a $1\,dm^3$ vessel. At eq. $[A] = 1.0$. Find $[B]$.

Further Practice （进阶）· 10

Further Practice

A. K_c expressions and simple values （K_c 表达式与代入）

1. Write K_c for $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$ and for $H_2 + I_2 \rightleftharpoons 2\,HI$.

2. At equilibrium $[H_2] = 0.20$, $[I_2] = 0.20$, $[HI] = 1.60\,mol\,dm^{-3}$. Calculate K_c for $H_2 + I_2 \rightleftharpoons 2\,HI$.

3. Write K_c for $2\,SO_2 + O_2 \rightleftharpoons 2\,SO_3$. If $[SO_2] = 0.20$, $[O_2] = 0.10$, $[SO_3] = 0.40$, find K_c.

B. Le Chatelier qualitative （勒夏特列定性）

4. For Haber $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$ ($\Delta H < 0$), state the effect of increasing pressure on the yield of NH_3.

5. For the same reaction, state the effect of increasing temperature on the yield of NH_3.

6. State the effect of a catalyst on the position of equilibrium and on K_c.

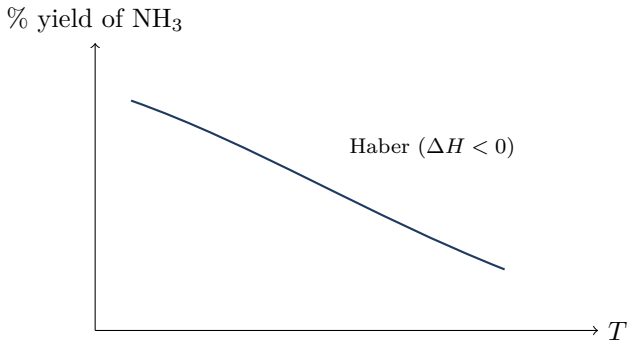
C. Acid–base basics （酸碱基础）

7. Identify acid and base in $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$ (Brønsted–Lowry).

8. Calculate $[H^+]$ in $0.0100\,\text{mol}\,\text{dm}^{-3}$ HCl (strong) and hence pH ($\text{pH} = -\log_{10}[H^+]$).

9. Write K_a for $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ and state what a small K_a indicates.

D. Equilibrium yield vs temperature (graph) （产率–温度图）



10. The graph shows equilibrium % yield of NH_3 vs temperature for Haber at fixed pressure.
(i) Explain why the yield falls as T rises.
(ii) Why is a moderate (not very low) temperature still used industrially?

Advanced Practice （培优）· 9

Advanced Practice

A. Equilibrium calculations from amounts （由起始量求 K_c）

1. $H_2 + CO_2 \rightleftharpoons H_2O + CO$ has $K_c = 1.0$. At equilibrium $[H_2] = [CO_2] = [H_2O] = 0.20\,\text{mol}\,\text{dm}^{-3}$. Find $[CO]$.

2. $0.200\,\text{mol}$ PCl_5 is placed in a $2.00\,\text{dm}^3$ vessel. At equilibrium $0.0400\,\text{mol}$ PCl_5 remains. Find

all equilibrium concentrations and K_c for $PCl_5 \rightleftharpoons PCl_3 + Cl_2$.

B. Conditions: yield vs rate vs K_c （产率 · 速率 · K_c 区分）

3. For Haber ($\Delta H < 0$), state the effect of $\uparrow T$, $\uparrow p$ and a catalyst on (i) equilibrium yield of NH_3 (ii) rate of reaching equilibrium.

4. Explain why K_c changes with temperature but does *not* change when concentration is altered at fixed T.

C. Coupled shifts and weak acids （多因素移动与弱酸）

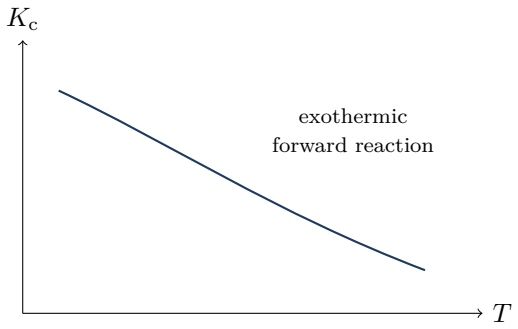
5. $N_2O_4 \rightleftharpoons 2\,NO_2$ is endothermic forward and increases the number of gas moles. Predict the effects of higher T and higher p on the equilibrium amount of NO_2, with brief reasons.

6. Two Haber plants are compared. Plant A uses 350°C with no catalyst. Plant B uses 450°C with an iron catalyst. For each plant, compare (i) equilibrium yield of NH_3, (ii) rate of reaching equilibrium, and (iii) likely daily production rate. Justify your choice of the more practical plant.

7. A weak acid HA is only partly dissociated. Explain why $[H^+]$ is much smaller than the formal concentration of HA, referring to K_a.

8. Adding OH^- to $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ at fixed T: predict the direction of shift and the effect on $[CH_3COOH]$.

D. K_c vs temperature (graph) （K_c–温度图）



9. The sketch shows how K_c varies with T for an exothermic forward reaction.
   (i) Explain the shape using Le Chatelier / van't Hoff ideas at AS level.
   (ii) At fixed T, concentration of a reactant is increased. Mark on the axes (or state) whether K_c changes, and what happens to the position of equilibrium.

8 Reaction Kinetics（反应动力学）

Collision theory; factors affecting rate; catalysts; Maxwell–Boltzmann distribution (qualitative).

Rules 规则

- **Rate of reaction**（反应速率）: change in concentration (or amount) of a reactant/product per unit time.
- **Collision theory**（碰撞理论）: particles must collide with sufficient energy (**activation energy**, E_a) and correct orientation.
- **Effective collision**（有效碰撞）: a collision that leads to reaction (energy $\geq E_a$ and correct orientation).
- **Concentration / pressure**（浓度/压强）: higher concentration (or gas pressure) $\Rightarrow$ more frequent collisions $\Rightarrow$ higher rate.
- **Temperature**（温度）: higher $T \Rightarrow$ higher average kinetic energy and a larger fraction of particles with $E \geq E_a \Rightarrow$ much higher rate.
- **Surface area**（表面积）: for solids, smaller particles / powder $\Rightarrow$ larger surface area $\Rightarrow$ higher rate.
- **Catalyst**（催化剂）: provides an alternative pathway with lower E_a; is not consumed overall.
- **Maxwell–Boltzmann**（麦克斯韦–玻尔兹曼分布）: curve of number of particles vs kinetic energy; shaded area beyond E_a = reacting fraction; catalyst lowers E_a and enlarges that area.

Worked Examples 例题

Example 1. State two conditions for an effective collision according to collision theory.

Sufficient energy ($\geq E_a$) and correct orientation.

Example 2. Explain, using collision theory, why increasing concentration of a reactant increases rate.

More particles per unit volume $\Rightarrow$ collisions occur more frequently $\Rightarrow$ more effective collisions per unit time $\Rightarrow$ higher rate.

Example 3. How does a catalyst increase rate? Refer to activation energy and the Maxwell–Boltzmann distribution.

Catalyst provides an alternative pathway with lower E_a.

On the Maxwell–Boltzmann curve, more particles have energy $\geq$ the lower E_a, so the fraction of effective collisions increases.

Basic Drill （入门综合）· 28

1. Define rate of reaction.

2. State what is meant by activation energy, E_a.

3. State two requirements for an effective collision.

4. Define a catalyst.

5. Does a catalyst appear in the overall chemical equation? Explain briefly.

6. State how rate usually changes when temperature is increased.

7. State how rate usually changes when reactant concentration is increased (homogeneous reaction).

8. For a reaction involving a solid, how does using a powder instead of large lumps affect rate?

9. State how increasing pressure affects the rate of a reaction between gases.

10. Sketch (describe) the Maxwell–Boltzmann distribution: axes labels and the meaning of the area under the curve.

11. On a Maxwell–Boltzmann curve, what does the area to the right of E_a represent?

12. Explain briefly why only a small fraction of collisions may be effective at room temperature.

13. Name one biological catalyst and state that it is a catalyst.

14. State whether a catalyst is used up in the reaction overall.

15. Give one industrial example where a catalyst is used (name reaction or process briefly).

16. Marble chips react with dilute HCl. Suggest one way to increase the rate without changing temperature or concentration.

17. Marble chips react with dilute HCl. Suggest one way to decrease the rate by changing concentration.

18. Explain why increasing temperature increases rate more than just “particles move faster”.

19. State the effect of a catalyst on ΔH of the reaction (qualitative).

20. State the effect of a catalyst on E_{a}.

21. Two Maxwell–Boltzmann curves are drawn at T_1 and T_2 ($T_2 > T_1$). How does the peak height and the high-energy tail change at T_2?

22. Why must colliding particles have the correct orientation?

23. In $\mathrm{CaCO_3(s) + 2\,HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)}$, name one measurable quantity that can indicate rate.

24. State one unit commonly used for rate of reaction involving concentration.

25. Distinguish between frequency of collisions and frequency of effective collisions.

26. Explain why finely divided iron is used in the Haber process rather than a single lump (rate focus).

27. A student says "a catalyst speeds up only the forward reaction". Correct this statement for a reversible reaction.

28. On one set of axes, describe how to show E_a with and without a catalyst on an energy profile (reaction pathway) diagram.

Practice （综合练习）· 12

Practice

1. Using collision theory, explain why increasing the pressure of a gaseous reactant mixture increases rate.

2. Explain, with reference to the Maxwell–Boltzmann distribution, why a small increase in temperature can cause a large increase in rate.

3. Compare the effect on rate of (i) doubling concentration of one aqueous reactant and (ii) adding a catalyst, in terms of collisions and E_a.

4. For $\mathrm{Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)}$, suggest two different experimental methods to follow rate and state what you would measure.

5. Explain why powdered $\mathrm{CaCO_3}$ reacts faster with acid than large chips of the same mass.

6. Draw / describe a Maxwell–Boltzmann curve and mark E_a. Shade the region corresponding to particles that can react.

7. A catalyst is added to a reaction. Describe two changes you would see on (i) the energy profile and (ii) the Maxwell–Boltzmann diagram.

8. Explain why orientation matters in a reaction between two asymmetric molecules (qualitative).

9. Zinc reacts with dilute sulfuric acid. Predict and explain the effect on rate of using more concentrated acid at the same temperature.

10. State and explain the effect on rate if the temperature is decreased, referring to both collision frequency and the energy distribution.

11. In a closed vessel, gases A and B react. The volume of the vessel is halved at constant temperature. Explain the effect on rate.

12. Distinguish a homogeneous catalyst from a heterogeneous catalyst and give one example of each type (name only).

Further Practice （进阶）· 9

Further Practice

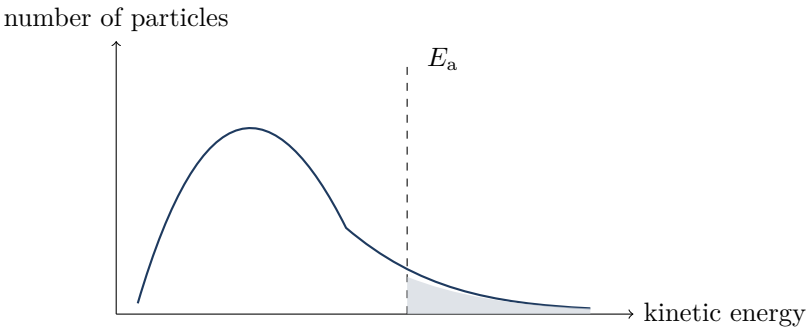
A. Collision theory factors （碰撞理论因素）

1. Using collision theory, explain why increasing the pressure of a gaseous reactant mixture increases rate.

2. Explain why powdered $CaCO_3$ reacts faster with acid than large chips of the same mass.

3. Zinc reacts with dilute sulfuric acid. Predict and explain the effect on rate of using more concentrated acid at the same temperature.

B. Maxwell–Boltzmann (with diagram) （麦克斯韦–玻尔兹曼图像）

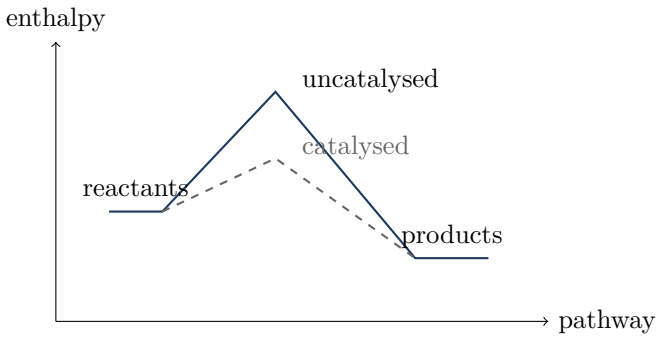


4. On the Maxwell–Boltzmann curve shown, what does the shaded area to the right of E_a represent?

5. Explain briefly why only a small fraction of collisions may be effective at room temperature.

6. Sketch (on the same axes idea) a second curve for a higher temperature $T_2 > T_1$. State how the peak height and the high-energy tail change at T_2.

C. Catalysts and energy profile （催化剂与能量剖面）



7. From the energy profiles: state the effect of a catalyst on E_a and on ΔH.

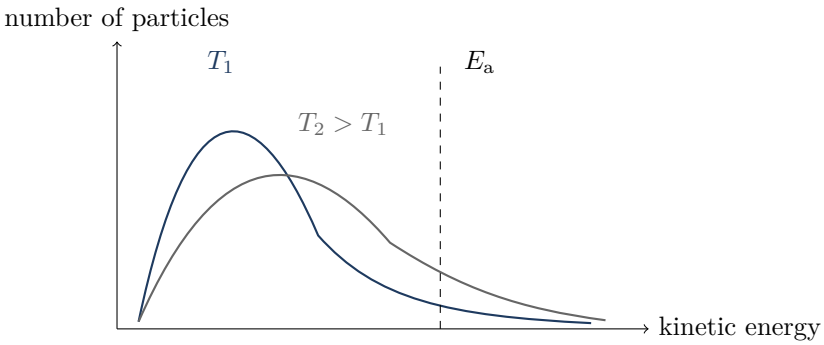
8. A student says "a catalyst speeds up only the forward reaction". Correct this statement for a reversible reaction.

9. Distinguish a homogeneous catalyst from a heterogeneous catalyst and give one example of each type (name only).

Advanced Practice （培优）· 8

Advanced Practice

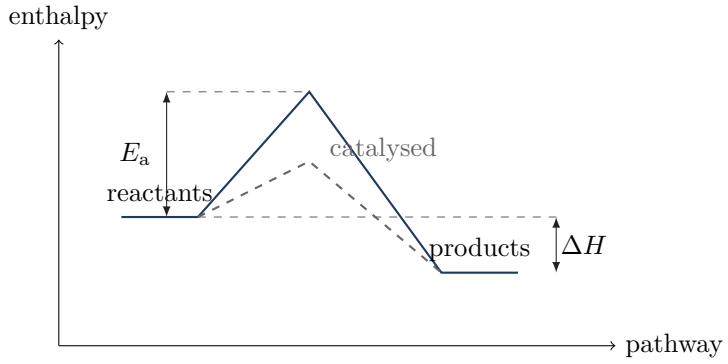
A. Maxwell–Boltzmann and temperature （分布曲线与温度）



1. The diagram shows Maxwell–Boltzmann curves at two temperatures for the same sample.
   (i) Which curve is T_2?
   (ii) Explain which curve has the larger fraction of particles able to react for the marked E_a.

2. Explain, with reference to the Maxwell–Boltzmann distribution, why a small increase in temperature can cause a large increase in rate.

B. Energy profiles and catalysts （能量剖面与催化）



3. On the energy profile, identify E_a (forward), E_a (reverse) and ΔH. State how a catalyst (dashed path) affects each of these three quantities.

4. Two pathways for the same reaction have similar collision frequency, but different activation energies: Pathway X: $E_a = 95\,\text{kJ mol}^{-1}$; Pathway Y: $E_a = 65\,\text{kJ mol}^{-1}$. Explain which pathway is faster and why the increase in successful-collision fraction is not linear with E_a.

C. Rate interpretation （速率变化解释）

5. For the reaction of marble with acid, the rate falls as the reaction proceeds even at constant temperature. Give two independent collision-theory reasons for this decrease, and propose one experimental modification that helps test which reason is dominant.

6. Explain why increasing temperature increases the proportion of successful collisions much more than it increases the total number of collisions per unit time.

7. A student claims that all collisions with energy greater than E_a lead to products. Evaluate this claim.

8. In a closed vessel, gases A and B react. The volume of the vessel is halved at constant temperature. Explain the effect on rate, referring to concentration and collisions.

9 Periodicity（元素周期性）

Period 3 trends; oxides and chlorides—structure, bonding, acid–base behaviour and electrical conductivity.

Rules 规则

- **Periodicity**（周期性）: repeating pattern of properties across Periods (e.g. Period 3: Na to Ar).
- **Atomic radius**（原子半径）: generally decreases Na → Cl (nuclear charge increases; electrons in same shell more strongly attracted).
- **Period 3 oxides**（第三周期氧化物）: Na_2O, MgO ionic; Al_2O_3 ionic with covalent character; SiO_2 giant covalent; P_4O_{10}, SO_2/SO_3, Cl_2O_7 molecular.
- **Acid–base of oxides**（氧化物酸碱性）: Na_2O, MgO basic; Al_2O_3 amphoteric; SiO_2 acidic (weakly); phosphorus, sulfur, chlorine oxides acidic.
  e.g. $Na_2O + H_2O \longrightarrow 2\,NaOH$; $SO_3 + H_2O \longrightarrow H_2SO_4$; $P_4O_{10} + 6\,H_2O \longrightarrow 4\,H_3PO_4$; amphoteric: $Al_2O_3 + 6\,HCl \longrightarrow 2\,AlCl_3 + 3\,H_2O$; $Al_2O_3 + 2\,NaOH + 3\,H_2O \longrightarrow 2\,NaAl(OH)_4$.
- **Period 3 chlorides**（第三周期氯化物）: NaCl, $MgCl_2$ ionic; $AlCl_3$ often treated as covalent / simple molecular (dimer Al_2Cl_6); $SiCl_4$, PCl_5/PCl_3, SCl_2 etc. molecular.
- **Hydrolysis / with water**（与水作用）: ionic chlorides dissolve (little/no hydrolysis for NaCl); $MgCl_2$ slight hydrolysis; $AlCl_3$, $SiCl_4$, PCl_5 react vigorously with water giving acidic solutions / HCl fumes.
  e.g. $SiCl_4 + 2\,H_2O \longrightarrow SiO_2 + 4\,HCl$; $PCl_5 + 4\,H_2O \longrightarrow H_3PO_4 + 5\,HCl$;
  $AlCl_3 + 3\,H_2O \longrightarrow Al(OH)_3 + 3\,HCl$ (or hexaaqua acid equation as taught).
- **Electrical conductivity**（导电性）: ionic oxides/chlorides conduct when molten (ions free); giant covalent and simple molecular do not (no free ions/electrons in the molten simple molecular case).
- **Structure link**（结构联系）: bonding & structure explain melting point trends, conductivity and acid–base behaviour across the Period.

Worked Examples 例题

Example 1. Classify Na_2O, Al_2O_3 and SO_3 as basic, acidic or amphoteric.

Na_2O basic; Al_2O_3 amphoteric; SO_3 acidic.

Example 2. State the structure and bonding in SiO_2 and whether molten SiO_2 conducts electricity.

Giant covalent structure (covalent bonding throughout).

Molten SiO_2 does not conduct (no free ions or delocalised electrons available for conduction in that sense).

Example 3. Describe what is observed when $SiCl_4$ is added to water, and write an equation.

Steamy fumes of HCl; acidic solution forms.

$SiCl_4 + 2\,H_2O \longrightarrow SiO_2 + 4\,HCl$
(or hydrated silica / silicic acid accepted at AS where consistent with syllabus equation).

Basic Drill （入门综合）· 28

1. Define periodicity.

2. State the trend in atomic radius across Period 3 from Na to Cl.

3. Explain briefly why atomic radius decreases across Period 3.

4. Write the formula of sodium oxide.

5. Write the formula of magnesium oxide.

6. Write the formula of aluminium oxide.

7. Write the formula of silicon dioxide.

8. Classify Na_2O as basic, acidic or amphoteric.

9. Classify MgO as basic, acidic or amphoteric.

10. Classify Al_2O_3 as basic, acidic or amphoteric.

11. Classify SiO_2 as basic, acidic or amphoteric.

12. Classify P_4O_{10} as basic, acidic or amphoteric.

13. Classify SO_3 as basic, acidic or amphoteric.

14. State the bonding and structure of Na_2O.

15. State the bonding and structure of SiO_2.

16. State the bonding and structure of SO_2.

17. Does molten NaCl conduct electricity? Explain briefly.

18. Does molten $SiCl_4$ conduct electricity? Explain briefly.

19. Write an equation for Na_2O reacting with water.

20. Write an equation for MgO reacting with dilute HCl.

21. Write an equation for SO_3 reacting with water.

22. Name the Period 3 chloride that is a giant ionic lattice: NaCl or $SiCl_4$?

23. State the approximate pH of the solution formed when NaCl dissolves in water.

24. State what is observed (fumes) when PCl_5 reacts with water.

25. Which oxide is amphoteric in Period 3?

26. Predict whether solid Na_2O conducts electricity at room temperature.

27. State the trend in character of Period 3 oxides from Na to S (basic $\rightarrow$ …).

28. Give the formula of the dimer commonly associated with aluminium chloride vapour.

Practice （综合练习）· 12

Practice

1. Compare the acid–base behaviour of Na_2O, Al_2O_3 and SO_3, giving one equation for each oxide with a suitable reagent (water or acid/base).

2. Explain why Na_2O has a high melting point whereas SO_2 has a low melting point.

3. Describe and explain the electrical conductivity of molten MgO and molten SiO_2.

4. Write equations for the reactions of Al_2O_3 with HCl(aq) and with hot concentrated NaOH(aq) (amphoteric behaviour).

5. Contrast the bonding in $MgCl_2$ and $SiCl_4$ and link this to their behaviour with water.

6. Explain why $SiCl_4$ fumes in moist air.

7. Across Period 3, melting points of the elements show a peak at Si. Explain in terms of structure.

8. State and explain whether aqueous solutions from NaCl and from $AlCl_3$ have the same pH.

9. Give the structures of P_4O_{10} and SiO_2 (giant or simple molecular) and relate to volatility.

10. Write an equation for Cl_2O_7 (or SO_3) with water to form an acid; name the acid formed for SO_3.

11. Predict and explain the conductivity of molten Al_2O_3 compared with molten SCl_2.

12. Summarise in one sentence how oxide structure changes from Na to Cl in Period 3.

Further Practice （进阶）· 10

Further Practice

A. Oxides: structure and acid–base （氧化物结构与酸碱）

1. Classify Na_2O, Al_2O_3 and SO_3 as basic, acidic or amphoteric.

2. State the bonding and structure of Na_2O, SiO_2 and SO_2.

3. Write equations for Na_2O with water and for SO_3 with water.

B. Chlorides and water （氯化物与水）

4. Does molten NaCl conduct electricity? Does molten $SiCl_4$? Explain briefly in each case.

5. Describe what is observed when $SiCl_4$ is added to water, and write an equation.

6. State and explain whether aqueous solutions from NaCl and from $AlCl_3$ have the same pH.

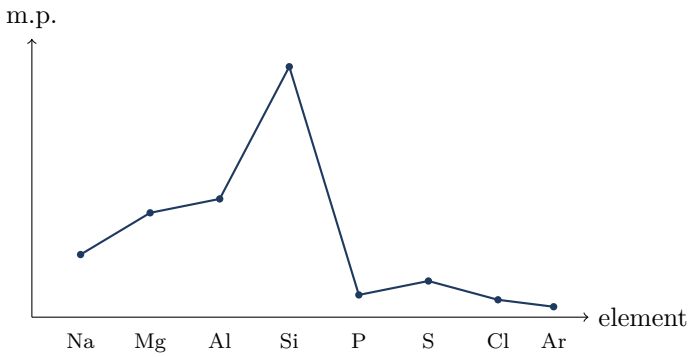
C. Trends across Period 3 （第三周期趋势识记）

7. State the trend in atomic radius across Period 3 from Na to Cl and explain briefly.

8. State the trend in character of Period 3 oxides from Na to S (basic $\rightarrow$ …).

9. Across Period 3, melting points of the elements show a peak at Si. Explain in terms of structure.

D. Melting point trend (graph) （熔点趋势图）



10. The sketch shows melting points of Period 3 elements (schematic).
    (i) Explain the high melting point of Si.
    (ii) Explain why P, S, Cl and Ar have much lower melting points than Si.

Advanced Practice （培优）· 8

Advanced Practice

A. Acid–base and bonding interpretation （酸碱性与成键解释）

1. Explain the trend in oxide acid–base character across Period 3 in terms of bonding and electronegativity / oxidation state (qualitative AS level).

2. Al_2O_3 is ionic with significant covalent character. How does this help explain why it is amphoteric and has a very high melting point?

B. Hydrolysis and covalent character （水解与共价成分）

3. Compare hydrolysis of NaCl, $MgCl_2$ and $SiCl_4$. Link observations to bonding and polarising power / covalent character.

4. Period 3 chlorides: arrange NaCl, $AlCl_3$, $SiCl_4$ in order of increasing covalent character and justify briefly.

C. Structure deduction （结构推断）

5. Using structure and bonding, explain why SiO_2 is a solid at room temperature but CO_2 is a gas, even though both involve Group 14 and oxygen.

6. A sample of a Period 3 chloride reacts vigorously with water giving steamy fumes and a strongly acidic solution; the molten chloride does not conduct. Identify the most likely type of chloride and suggest a formula.

7. Describe and explain the electrical conductivity of molten MgO and molten SiO_2.

8. Write equations for the reactions of Al_2O_3 with HCl(aq) and with hot concentrated NaOH(aq) (amphoteric behaviour).

10 Group 2, Group 17, N & S（族化学）

Group 2 trends and key reactions; halogen displacement and redox; simple nitrogen and sulfur chemistry.

Rules 规则

- **Group 2 trends**（第 2 族趋势）: atomic/ionic radius increases down the group; ionisation energy decreases; reactivity with water generally increases Mg → Ba.
- **Group 2 with water / acid**（与水/酸）: Mg reacts slowly with cold water, more readily with steam; Ca, Sr, Ba react with cold water giving hydroxide + H_2; all react with dilute acids giving salt + H_2.
  e.g. $Ca + 2\,H_2O \longrightarrow Ca(OH)_2 + H_2$; $Mg + H_2O \longrightarrow MgO + H_2$ (steam);
  $Mg + 2\,HCl \longrightarrow MgCl_2 + H_2$.
- **Group 2 oxides / hydroxides**（氧化物/氢氧化物）: basic; solubility of hydroxides increases down the group; solubility of sulfates decreases down the group.
  e.g. $CaO + H_2O \longrightarrow Ca(OH)_2$; $Ca(OH)_2 + 2\,HCl \longrightarrow CaCl_2 + 2\,H_2O$.
- **Thermal stability**（热稳定性）: Group 2 carbonates and nitrates become more thermally stable down the group (larger cation, less polarising).
  $MCO_3 \longrightarrow MO + CO_2$; $2\,M(NO_3)_2 \longrightarrow 2\,MO + 4\,NO_2 + O_2$.
- **Group 17 displacement**（卤素置换）: a more reactive halogen displaces a less reactive halide from solution; reactivity $F_2 > Cl_2 > Br_2 > I_2$.
  e.g. $Cl_2 + 2\,Br^- \longrightarrow 2\,Cl^- + Br_2$; $Cl_2 + 2\,I^- \longrightarrow 2\,Cl^- + I_2$; $Br_2 + 2\,I^- \longrightarrow 2\,Br^- + I_2$.
- **Halogen redox**（卤素氧化还原）: halogens are oxidising agents; oxidising power decreases down the group; halide reducing power increases down the group.
  e.g. $Cl_2 + 2\,e^- \longrightarrow 2\,Cl^-$ (reduction half-equation).
- **Nitrogen**（氮）: N_2 very unreactive (strong triple bond); Haber: $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$;
  ammonia is a weak base: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$;
  oxidation (AS awareness): $4\,NH_3 + 5\,O_2 \longrightarrow 4\,NO + 6\,H_2O$; $2\,NO + O_2 \longrightarrow 2\,NO_2$.
- **Sulfur**（硫）: SO_2 from burning S or roasting sulfide ores; Contact process idea; acidic solutions; acid rain link.
  $S + O_2 \longrightarrow SO_2$; $2\,SO_2 + O_2 \rightleftharpoons 2\,SO_3$; $SO_2 + H_2O \longrightarrow H_2SO_3$; $SO_3 + H_2O \longrightarrow H_2SO_4$.

Worked Examples 例题

Example 1. Write an equation for calcium reacting with cold water.

$Ca + 2\,H_2O \longrightarrow Ca(OH)_2 + H_2$

Example 2. Chlorine water is added to aqueous potassium bromide. State the observation and write an ionic equation.

Colourless/pale solution turns orange/brown (bromine formed).

$Cl_2 + 2\,Br^- \longrightarrow 2\,Cl^- + Br_2$

Example 3. State the trend in solubility of Group 2 hydroxides down the group, and give the formula of the hydroxide of barium.

Solubility increases down the group.

$Ba(OH)_2$

Basic Drill （入门综合）· 30

1. State the trend in atomic radius down Group 2.

2. State the trend in first ionisation energy down Group 2.

3. Write the electron configuration of Mg^{2+} in terms of shells or noble-gas shorthand if used in your course.

4. Write an equation for magnesium burning in oxygen.

5. Write an equation for calcium reacting with cold water.

6. Write an equation for magnesium reacting with steam.

7. Write an equation for Mg with dilute HCl.

8. State the trend in solubility of Group 2 hydroxides down the group.

9. State the trend in solubility of Group 2 sulfates down the group.

10. Which Group 2 hydroxide is sparingly soluble: $Mg(OH)_2$ or $Ba(OH)_2$?

11. Write an equation for the thermal decomposition of $CaCO_3$.

12. Write an equation for the thermal decomposition of $Mg(NO_3)_2$.

13. State the trend in thermal stability of Group 2 carbonates down the group.

14. List the halogens in order of decreasing reactivity: Cl_2, Br_2, I_2.

15. Write an ionic equation for Cl_2 displacing bromide ions.

16. Write an ionic equation for Cl_2 displacing iodide ions.

17. Write an ionic equation for Br_2 displacing iodide ions.

18. State the colour of aqueous bromine (approximate).

19. State the colour of aqueous iodine (approximate, in water or with iodide).

20. Does I_2 displace Cl^- from solution? Explain briefly.

21. State the trend in oxidising power of the halogens down Group 17.

22. State the trend in reducing power of the halide ions down Group 17.

23. Write the equation for the Haber process (include equilibrium arrow).

24. Why is N_2 relatively unreactive at room temperature?

25. Write an equation for ammonia acting as a base with H^+ (or with water).

26. Write an equation for sulfur burning in air to form SO_2.

27. Write an equation for SO_2 dissolving/reacting with water (formation of acid).

28. Write an equation for SO_3 with water.

29. Name the industrial process that oxidises SO_2 to SO_3 (catalyst name optional).

30. Give one environmental problem associated with SO_2 emissions.

Practice （综合练习）· 14

Practice

1. Explain why reactivity of Group 2 metals with water increases down the group.

2. Compare observations when Mg and Ca are added separately to cold water.

3. Explain the trend in thermal stability of Group 2 carbonates down the group in terms of cation size and polarising power.

4. Aqueous chlorine is mixed with aqueous potassium iodide. State colour change and write the ionic equation.

5. Explain, using redox ideas, why chlorine can oxidise bromide but iodine cannot oxidise chloride.

6. Write equations for the reactions of Ba with water and of $Ba(OH)_2$ with dilute HCl.

7. State and explain which is more soluble: $MgSO_4$ or $BaSO_4$.

8. Describe a simple test-tube reaction to show that chlorine is a stronger oxidising agent than bromine (displacement).

9. Write equations for the Contact process steps: S to SO_2; SO_2 to SO_3; SO_3 to sulfuric acid (via oleum route or direct hydration as in your syllabus).

10. Explain why the Haber process uses a catalyst and is typically run at moderate temperature (qualitative conflict of rate vs yield).

11. Ammonia solution turns red litmus blue. Write an equation to explain this alkaline behaviour.

12. Write an equation for the formation of NO from N_2 and O_2 at high temperature.

13. Group 2 nitrates decompose on heating. Write a general equation for $M(NO_3)_2$ and state how ease of decomposition changes down the group.

14. Predict the products when chlorine water is added to aqueous sodium bromide, and state how you would observe bromine has formed.

Further Practice （进阶）· 9

Further Practice

A. Group 2 trends and reactions （第 2 族趋势与反应）

1. State the trends down Group 2 in (i) atomic radius (ii) first ionisation energy.

2. Write equations for Ca with cold water and for Mg with steam.

3. State the trends in solubility of Group 2 hydroxides and of Group 2 sulfates down the group.

B. Halogen displacement （卤素置换）

4. Write ionic equations for Cl_2 with aqueous bromide and with aqueous iodide.

5. State observations when aqueous chlorine is mixed with aqueous potassium iodide.

6. Explain, using redox ideas, why chlorine can oxidise bromide but iodine cannot oxidise chloride.

C. Nitrogen and sulfur basics （氮与硫基础）

7. Write the Haber equation and state why N_2 is relatively unreactive at room temperature.

8. Write equations for sulfur burning in air and for SO_3 with water.

9. Give one environmental problem associated with SO_2 emissions.

Advanced Practice （培优）· 8

Advanced Practice

A. Solubility and thermal stability （溶解度与热稳定性）

1. Explain the trends in solubility of Group 2 hydroxides and sulfates down the group (state trends clearly; brief lattice/hydration energy idea if required by your course).

2. Explain the trend in thermal stability of Group 2 carbonates down the group in terms of cation size and polarising power.

B. Redox and reactivity comparisons （氧化还原与反应性对比）

3. Construct a redox half-equation approach: show that in $Cl_2 + 2\,Br^- \longrightarrow 2\,Cl^- + Br_2$, chlorine is reduced and bromide is oxidised.

4. Compare the reactions of Mg and Ba with water and relate to ionisation energy / ease of

forming M^{2+}.

C. Acid rain and decomposition patterns （酸雨与分解规律）

5. SO_2 and NO_2 both contribute to acid rain. Write equations showing how each can lead to acid in the environment (simple AS equations).

6. Thermal decomposition: contrast products from heating $Ca(NO_3)_2$ with heating $NaNO_3$ (Group 1 vs Group 2 pattern awareness at AS).

7. Evaluate why F_2 is not used in school-style aqueous displacement tests alongside $Cl_2/Br_2/I_2$. Give two chemical reasons and one consequence for practical experiment design.

8. Explain why the Haber process uses a catalyst and is typically run at moderate temperature (qualitative conflict of rate vs yield).

11 Organic Foundations（有机入门）

Homologous series; naming alkanes, alkenes and alcohols (simple); isomers; functional groups.

Rules 规则

- **Homologous series**（同系物系列）: same general formula and functional group; each member differs by CH_2; gradual trend in physical properties.
- **General formulas**（通式）: alkanes C_nH_{2n+2}; alkenes (one C=C) C_nH_{2n}; alcohols (one OH) $C_nH_{2n+1}OH$.
- **Functional group**（官能团）: the reactive group that largely determines chemistry, e.g. C=C, −OH, −COOH, halogeno −X.
- **Naming (IUPAC simple)**（简单命名）: longest chain; number to give lowest locants; prefixes methyl/ethyl; endings -ane, -ene, -ol.
- **Structural isomers**（结构异构体）: same molecular formula, different structural formula (chain, position, functional-group isomers).
- **Displayed / structural / skeletal**（表示法）: displayed shows all atoms/bonds; structural may be condensed; skeletal shows carbon framework.
- **Saturated vs unsaturated**（饱和与不饱和）: alkanes saturated (only single C–C); alkenes unsaturated (contain C=C). A quick test is decolourisation of bromine water by addition across C=C:

  $CH_2{=}CH_2 + Br_2 \longrightarrow CH_2BrCH_2Br$.
- **Empirical / molecular**（经验式/分子式）: still apply in organic: e.g. ethene molecular C_2H_4, empirical CH_2.
- **Isomer checklist**（异构体检查）: keep the molecular formula unchanged; alter only connectivity; then check longest chain, functional-group position and whether two drawings are actually the same structure after rotation.

Worked Examples 例题

Example 1. Give the general formula of alkanes and name $CH_3CH_2CH_3$.

Alkanes: C_nH_{2n+2}.

Name: propane.

Example 2. Name $CH_3CH(OH)CH_3$ and state its functional group.

Propan-2-ol.

Functional group: hydroxyl (−OH), alcohol.

Example 3. Draw/describe two structural isomers of C_4H_{10} and name them.

Butane (straight chain) and methylpropane (branched).

Both have molecular formula C_4H_{10}.

Basic Drill （入门综合）· 32

1. Define a homologous series.

2. State the general formula of the alkanes.

3. State the general formula of the alkenes (one double bond).

4. State the general formula of the alcohols (one $-OH$).

5. Define a functional group.

6. Name CH_4.

7. Name CH_3CH_3.

8. Name $CH_3CH_2CH_3$.

9. Name $CH_3CH_2CH_2CH_3$.

10. Name $CH_2{=}CH_2$.

11. Name $CH_3CH{=}CH_2$.

12. Name CH_3OH.

13. Name CH_3CH_2OH.

14. Name $CH_3CH_2CH_2OH$.

15. Name $CH_3CH(OH)CH_3$.

16. What is the functional group in an alkene?

17. What is the functional group in an alcohol?

18. State whether propane is saturated or unsaturated.

19. State whether propene is saturated or unsaturated.

20. Give the molecular formula of butane.

21. Give the molecular formula of butene (one C=C).

22. Define structural isomers.

23. Name the straight-chain and branched isomers of C_4H_{10}.

24. Give the displayed formula description of ethane (state all C and H with single bonds).

25. What is the next member after ethane in the alkane homologous series? Give formula and name.

26. By what group do successive alkanes differ?

27. Identify the homologous series of CH_3CH_2COOH (name the series).

28. Identify the functional group in CH_3CHO (aldehyde).

29. Identify the functional group in CH_3COCH_3 (ketone).

30. Name $(CH_3)_3CBr$ as a halogenoalkane (simple: 2-bromo-2-methylpropane).

31. State one physical property that generally increases down an alkane homologous series.

32. Give the empirical formula of ethene C_2H_4.

Practice （综合练习）· 14

Practice

1. Explain why alkanes form a homologous series, referring to formula and properties.

2. Name the following: $CH_3CH_2CH(CH_3)CH_3$; $CH_2{=}CHCH_2CH_3$; $CH_3CH_2CH_2CH_2OH$.

3. Draw condensed structural formulas for all structural isomers of C_3H_8O that are alcohols, and name them.

4. Distinguish chain isomerism and position isomerism, each with one pair of examples (names or formulas).

5. Explain why C_2H_6 and C_3H_8 have similar chemical properties but different boiling points.

6. Give the skeletal formula description of butane and methylpropane (words: unbranched vs one methyl branch on C2 of a C3 chain).

7. Classify each as alkane, alkene or alcohol: C_3H_6, C_3H_8, C_3H_7OH.

8. Name $CH_3CH{=}CHCH_3$ including the need for a number, and explain why numbering is required.

9. State the functional group and homologous series for: (i) CH_3COOH (ii) CH_3NH_2 (iii) CH_3Cl.

10. Write molecular formulas for the first four alkenes and name them.

11. Isomers of C_4H_8 include alkenes. Name but-1-ene and but-2-ene and state the type of isomerism.

12. Give the structural formula of 2-methylpropan-1-ol and confirm it fits C_4H_9OH.

13. Define unsaturated hydrocarbon and give one chemical test idea that distinguishes ethene from ethane (decolourising bromine water—observation only).

14. A hydrocarbon has molecular formula C_5H_{12}. Deduce whether it can be an alkene; explain.

Further Practice （进阶）· 9

Further Practice

A. Systematic naming and formulae （系统命名与分子式）

1. Name $CH_3CH(CH_3)CH_2CH_3$ and state the molecular formula.

2. Write a condensed structural formula for 3-methylpent-1-ene.

3. Name $CH_3CH_2CH(OH)CH_3$, giving the lowest possible locant to the principal functional group.

B. Structural isomerism （结构异构分类与列举）

4. Draw and name all alcohol structural isomers with molecular formula C_3H_8O.

5. Give one pair of chain isomers and one pair of position isomers with molecular formula C_4H_8.

6. Ethanol and methoxymethane both have formula C_2H_6O. State the type of isomerism and identify the functional group in each.

C. Representation and classification （结构表示与官能团判断）

7. Convert the name 2-methylpropan-1-ol into a condensed structural formula and classify the alcohol as primary, secondary or tertiary.

8. A compound has the skeletal structure of a four-carbon chain with a C=C bond beginning at C2. Give its name and molecular formula.

9. Compounds A and B are consecutive members of the same homologous series. A is C_3H_6; B has one additional CH_2 unit. Give the formula and one possible name of B.

Advanced Practice （培优）· 8

Advanced Practice

A. Isomer deduction under constraints （限制条件下的异构体推断）

1. List all four alcohol structural isomers of $C_4H_{10}O$; name and classify each as primary, secondary or tertiary.

2. A compound has formula C_4H_8, contains one C=C bond and has no E/Z isomerism. Give the two possible structural formulas and names.

B. Naming from complex structures （复杂结构的母链与编号）

3. Name $CH_2{=}C(CH_3)CH_2CH_3$. Explain how the parent chain and numbering are selected.

4. Draw a condensed structural formula for 3,3-dimethylbutan-2-ol and explain why the parent chain must include the carbon bearing –OH.

C. Structure–property reasoning （结构与物理性质关联）

5. Put pentane, 2-methylbutane and 2,2-dimethylpropane in decreasing boiling-point order and explain the effect of branching.

6. Ethanol and methoxymethane have the same molecular formula C_2H_6O. Predict which has the higher boiling point and greater solubility in water, and explain both predictions using intermolecular forces.

D. Formula and identity from combined data （通式与信息综合）

7. Hydrocarbon X contains 85.7% carbon by mass and has $M_r = 56$. Determine its empirical and molecular formula, then state two possible homologous-series types.

8. Compounds P and Q both have molecular formula C_3H_6O. P contains a terminal –CHO group; Q contains a carbonyl carbon bonded to two carbon atoms. Name P and Q and state their relationship.

12 Hydrocarbons & Halogen Compounds（烃与卤代烃）

Combustion; addition to alkenes; free-radical halogenation (outline); nucleophilic substitution of halogenoalkanes (NaOH aq / ethanolic).

Rules 规则

- **Combustion**（燃烧）: alkanes/alkenes burn in excess O_2 to CO_2 + H_2O; limited O_2 may give CO or C (soot).
  Complete: $C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O$;
  incomplete to CO: $2\,C_3H_8 + 7\,O_2 \longrightarrow 6\,CO + 8\,H_2O$.
- **Addition to alkenes**（烯烃加成）: C=C undergoes addition; the π part of C=C breaks and two new single bonds form:
  hydrogenation: $CH_2{=}CH_2 + H_2 \longrightarrow CH_3CH_3$ (Ni, heat);
  hydrogen halide: $CH_2{=}CH_2 + HBr \longrightarrow CH_3CH_2Br$;
  bromination: $CH_2{=}CH_2 + Br_2 \longrightarrow CH_2BrCH_2Br$ (orange/brown → colourless);
  hydration: $CH_2{=}CH_2 + H_2O \longrightarrow CH_3CH_2OH$ (steam, H_3PO_4, heat/pressure).
- **Electrophilic addition idea**（亲电加成概念）: double bond attacked; product is saturated addition compound.
- **Free-radical substitution**（自由基取代 · 概要）: alkanes + X_2 (UV): initiation, propagation and termination; product halogenoalkane (+ HX).
  Overall: $CH_4 + Cl_2 \longrightarrow CH_3Cl + HCl$ (UV);
  initiation: $Cl_2 \longrightarrow 2\,Cl\cdot$;
  propagation: $Cl\cdot + CH_4 \longrightarrow HCl + CH_3\cdot$, $CH_3\cdot + Cl_2 \longrightarrow CH_3Cl + Cl\cdot$;
  termination examples: $Cl\cdot + Cl\cdot \longrightarrow Cl_2$, $CH_3\cdot + Cl\cdot \longrightarrow CH_3Cl$, $CH_3\cdot + CH_3\cdot \longrightarrow C_2H_6$.
- **Halogenoalkanes**（卤代烷）: contain C–X; classified primary / secondary / tertiary.
- **Nucleophilic substitution**（亲核取代）: aqueous NaOH (heat): $-OH$ replaces X to form an alcohol.
  $CH_3CH_2Br + OH^- \longrightarrow CH_3CH_2OH + Br^-$.
- **Elimination with ethanolic NaOH**（醇钠碱消除）: hot ethanolic NaOH: H and X are removed from adjacent carbon atoms to form C=C.
  $CH_3CH_2Br + OH^- \longrightarrow CH_2{=}CH_2 + H_2O + Br^-$.
- **Reagent conditions matter**（条件重要）: same NaOH: aqueous ⇒ substitution; ethanolic / heat ⇒ elimination (AS contrast).

Worked Examples 例题

Example 1. Write a balanced equation for the complete combustion of propane.

$C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O$

Example 2. State the observation when ethene is bubbled through bromine water, and write an equation.

Bromine water decolourises (orange/brown → colourless).

$CH_2{=}CH_2 + Br_2 \longrightarrow CH_2BrCH_2Br$

Example 3. Give reagents and conditions to convert bromoethane to ethanol, and write an equation.

Heat with aqueous NaOH (or KOH).

$CH_3CH_2Br + NaOH \longrightarrow CH_3CH_2OH + NaBr$
(or $CH_3CH_2Br + OH^- \longrightarrow CH_3CH_2OH + Br^-$).

Basic Drill （入门综合）· 30

1. Write the equation for complete combustion of methane.

2. Write the equation for complete combustion of ethene.

3. State two products of incomplete combustion of a hydrocarbon.

4. What type of reaction occurs when Br_2 adds across C=C?

5. Write an equation for ethene reacting with Br_2.

6. Write an equation for ethene reacting with HBr.

7. Write an equation for ethene reacting with H_2 (name a catalyst).

8. Write an equation for ethene reacting with steam to form ethanol (acid catalyst).

9. State the colour change when bromine water reacts with an alkene.

10. Does ethane decolourise bromine water in the dark under normal lab conditions? (Expect no.)

11. Name the mechanism type for halogenation of methane in UV light (free-radical substitution).

12. Write an equation for the overall reaction of methane with chlorine (monosubstitution).

13. In free-radical chlorination, what is meant by initiation?

14. Give one example of a propagation step for $CH_4 + Cl_2$ (UV).

15. Give one example of a termination step in chlorination of methane.

16. Why is UV light needed for the reaction of CH_4 with Cl_2?

17. Define a nucleophile (simple: electron-pair donor / attacks $\delta+$ carbon).

18. Classify CH_3CH_2Br as primary, secondary or tertiary halogenoalkane.

19. Classify $(CH_3)_3CBr$ as primary, secondary or tertiary.

20. Give reagents/conditions for converting bromoethane to ethanol.

21. Write an equation for bromoethane with aqueous NaOH.

22. Give reagents/conditions for converting bromoethane to ethene.

23. Write an equation for bromoethane with ethanolic NaOH (elimination).

24. State the role of aqueous vs ethanolic solvent for NaOH with a halogenoalkane.

25. Name the product when propene reacts with HBr (major product awareness: 2-bromopropane).

26. Write an equation for propene + Br_2.

27. What is formed when ethene undergoes addition polymerisation? (poly(ethene)—name only).

28. State whether combustion of alkanes is addition, substitution or redox/combustion classification.

29. Give the structural formula of the addition product of ethene and H_2O (ethanol).

30. Write the ionic equation for OH^- substituting in bromoethane.

Practice （综合练习）· 14

Practice

1. Balance and compare complete combustion equations for ethane and ethene; which needs more O_2 per mole of hydrocarbon?

2. Explain why alkenes decolourise bromine water but alkanes do not (under usual conditions).

3. Outline the free-radical substitution of ethane with chlorine: write initiation, one propagation pair, and one termination.

4. Give reagents to convert ethene into: (i) ethane (ii) 1,2-dibromoethane (iii) bromoethane.

5. Explain why UV light is essential in methane chlorination but not for ethene + Br_2 (aq).

6. Contrast the organic products when 1-bromopropane is heated with (i) aqueous NaOH (ii) ethanolic NaOH.

7. Write equations for both reactions in Q6.

8. Propene reacts with HBr. Name both possible addition products and state which is the major product (Markownikoff, simple).

9. Classify the reactions: ethene + H_2; methane + Cl_2 (UV); bromoethane + aqueous NaOH (addition / substitution types).

10. Explain what is meant by homolytic fission in the initiation step of chlorination.

11. Why can multiple substitution products form when methane reacts with excess chlorine?

12. Write an equation for complete combustion of C_4H_{10} and calculate (qualitative) whether more or fewer moles of CO_2 form than from 1 mol methane.

13. Give the mechanism name for OH^- attacking bromoethane, and state what bond breaks.

14. Design a two-step sequence ethene → bromoethane → ethanol (reagents only).

Further Practice （进阶）· 9

Further Practice

A. Alkene addition reactions （烯烃加成与条件）

1. Write the equation, reagent and condition for converting ethene to ethane.

2. Propene reacts with bromine. Write the structural formula and name of the organic product, and state the observation.

3. State the reagent and conditions for hydration of ethene, and write the equation for formation of ethanol.

B. Free-radical substitution （自由基链反应三阶段）

4. Write the initiation step in the chlorination of methane and name the type of bond fission.

5. Write the two propagation steps that form chloromethane from methane.

6. Give two different termination equations and explain why each terminates the chain.

C. Substitution versus elimination （水解与消除条件对比）

7. Write the ionic equation for bromoethane heated with aqueous hydroxide ions and name the mechanism.

8. Write the equation for 2-bromopropane heated with ethanolic KOH; name the organic product.

9. A student heats 1-bromobutane separately with aqueous and ethanolic KOH. State the principal organic product in each experiment.

Advanced Practice （培优）· 8

Advanced Practice

A. Product selectivity in addition and elimination （加成与消除的产物选择）

1. Propene reacts with HBr. Draw both possible structural products and explain why 2-bromopropane is the major product.

2. 2-bromobutane is heated with ethanolic KOH. Name the possible alkene products and identify which product can show E/Z isomerism.

B. Mechanism evidence and comparison （机理证据与反应对比）

3. Compare ethane reacting with Br_2 under UV light and ethene reacting with bromine water. State mechanism, conditions, observation and organic product for each.

4. In methane chlorination, explain why traces of ethane can form even when ethane was not initially present. Include the radical equation.

C. Multi-step organic routes （多步路线与条件选择）

5. Starting from ethane, devise a two-step route to ethanol through a halogenoalkane. Give reagents, conditions and equations.

6. Starting from 2-bromopropane, outline a two-step route to propan-2-ol through an alkene. Give the reagent and condition for each step.

D. Stoichiometry and environmental chemistry （燃烧计量与环境影响）

7. A $10.0\,\mathrm{cm}^3$ gaseous hydrocarbon requires $50.0\,\mathrm{cm}^3$ O_2 for complete combustion and forms $30.0\,\mathrm{cm}^3$ CO_2, all volumes measured under the same conditions. Deduce the molecular formula.

8. Write equations for complete combustion and incomplete combustion to CO of propane. Explain why the second process presents a greater immediate health risk.

13 Oxygen & Nitrogen Organics + Synthesis（含氧含氮与合成）

Alcohols, carbonyls, carboxylic acids and simple derivatives; amines intro; addition polymerisation; short two-step synthesis.

Rules 规则

- **Alcohol classification**（醇的分类）: primary, secondary, tertiary (number of C atoms attached to the C bearing −OH).
- **Oxidation of alcohols**（醇的氧化）: acidified dichromate(VI) is used; distil an aldehyde as it forms, but use excess oxidant under reflux for an acid.
  primary to aldehyde: $\mathrm{CH_3CH_2OH + [O] \longrightarrow CH_3CHO + H_2O}$;
  primary to acid: $\mathrm{CH_3CH_2OH + 2\,[O] \longrightarrow CH_3COOH + H_2O}$;
  secondary to ketone: $\mathrm{CH_3CH(OH)CH_3 + [O] \longrightarrow CH_3COCH_3 + H_2O}$;
  tertiary alcohols are not readily oxidised under these conditions.
- **Carbonyls**（羰基化合物）: aldehydes RCHO and ketones RCOR' both contain C=O. Tollens'/Fehling's oxidise an aldehyde (positive) but not a ketone; the organic change is:
  $\mathrm{RCHO + [O] \longrightarrow RCOOH}$.
- **Carboxylic acids**（羧酸）: −COOH; weak acids; react with metals, carbonates and bases; form esters with alcohols.
  metal: $\mathrm{2\,CH_3COOH + Mg \longrightarrow (CH_3COO)_2Mg + H_2}$;
  carbonate: $\mathrm{2\,CH_3COOH + Na_2CO_3 \longrightarrow 2\,CH_3COONa + H_2O + CO_2}$;
  base: $\mathrm{CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O}$;
  esterification: $\mathrm{CH_3COOH + CH_3CH_2OH \rightleftharpoons CH_3COOCH_2CH_3 + H_2O}$ (conc. $\mathrm{H_2SO_4}$, heat).
- **Esters & acyl chlorides (simple)**（酯与酰氯 · 基础）: acyl chlorides react readily with alcohols to give esters, and with ammonia to give amides.
  $\mathrm{CH_3COCl + CH_3CH_2OH \longrightarrow CH_3COOCH_2CH_3 + HCl}$;
  $\mathrm{CH_3COCl + 2\,NH_3 \longrightarrow CH_3CONH_2 + NH_4Cl}$.
- **Amines (intro)**（胺 · 入门）: $\mathrm{-NH_2}$ (primary amine); weak bases; accept $\mathrm{H^+}$ and form salts with acids.
  $\mathrm{CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-}$; $\mathrm{CH_3NH_2 + HCl \longrightarrow CH_3NH_3^+Cl^-}$.
- **Addition polymerisation**（加成聚合）: alkene C=C opens and monomers join without elimination of a small molecule.
  $\mathrm{n\,CH_2{=}CH_2 \longrightarrow [-CH_2-CH_2-]_n}$.
- **Two-step synthesis**（两步合成）: plan reagents/conditions for each step; common links include:
  $\mathrm{CH_2{=}CH_2 + H_2O \longrightarrow CH_3CH_2OH}$ (steam/$\mathrm{H^+}$); $\mathrm{CH_3CH_2Br + OH^- \longrightarrow}$

$CH_3CH_2OH + Br^-$;
then use the oxidation and esterification equations above as required.

Worked Examples 例题

Example 1. Classify CH_3CH_2OH, $CH_3CH(OH)CH_3$ and $(CH_3)_3COH$ as primary, secondary or tertiary.

CH_3CH_2OH primary; $CH_3CH(OH)CH_3$ secondary; $(CH_3)_3COH$ tertiary.

Example 2. State reagents/conditions to oxidise ethanol to ethanoic acid, and write an equation using [O].

Heat under reflux with acidified potassium dichromate(VI) (excess oxidant).

$CH_3CH_2OH + 2\,[O] \longrightarrow CH_3COOH + H_2O$

Example 3. Write an equation for ethanoic acid with ethanol (conc. H_2SO_4) and name the ester.

$CH_3COOH + CH_3CH_2OH \rightleftharpoons CH_3COOCH_2CH_3 + H_2O$

Ethyl ethanoate.

Basic Drill （入门综合）· 32

1. Define a primary alcohol.

2. Define a secondary alcohol.

3. Define a tertiary alcohol.

4. Classify CH_3OH as primary, secondary or tertiary.

5. Classify $CH_3CH(OH)CH_3$ as primary, secondary or tertiary.

6. Classify $(CH_3)_3COH$ as primary, secondary or tertiary.

7. Name the organic product when ethanol is oxidised carefully (distil) with acidified dichromate.

8. Name the organic product when ethanol is fully oxidised under reflux with excess acidified dichromate.

9. Write an [O] equation for ethanol $\rightarrow$ ethanal.

10. Write an [O] equation for ethanol $\rightarrow$ ethanoic acid.

11. Name the product of oxidation of propan-2-ol.

12. Are tertiary alcohols readily oxidised by acidified dichromate?

13. What functional group is present in ethanal?

14. What functional group is present in propanone?

15. State the observation when Tollens' reagent is warmed with an aldehyde.

16. State the result of Tollens' reagent with a ketone.

17. Write an equation for ethanoic acid with sodium carbonate (ions or molecular).

18. Write an equation for ethanoic acid with NaOH.

19. Write an equation for esterification of ethanoic acid with methanol.

20. Name CH_3COOCH_3.

21. Name $CH_3COOCH_2CH_3$.

22. What gas is evolved when a carboxylic acid reacts with a carbonate?

23. Give the general structure of a primary amine.

24. Write an equation for CH_3NH_2 with HCl.

25. Is methylamine acidic, basic or neutral in water? (basic)

26. Name the polymer formed from ethene.

27. Draw/describe the repeating unit of poly(ethene).

28. Name the polymer formed from chloroethene (PVC).

29. Give reagents for ethanol → ethene (dehydration).

30. Give reagents for ethene → ethanol (hydration).

31. Outline two steps: ethanol → ethanoic acid (one transformation may count as oxidation under reflux).

32. Outline two steps to convert ethene to ethanol via bromoethane.

Practice （综合练习）· 14

Practice

1. Explain how to obtain ethanal rather than ethanoic acid from ethanol using acidified dichromate (conditions).

2. Write [O] equations for (i) propan-1-ol → propanal (ii) propan-1-ol → propanoic acid (iii) propan-2-ol → propanone.

3. Describe chemical tests to distinguish ethanal from propanone (reagent and observation for each).

4. Write equations for ethanoic acid with (i) Mg (ii) Na_2CO_3 (iii) CH_3CH_2OH/H^+.

5. Give reagents/conditions to make ethyl ethanoate from ethanol in two different routes (acid + alcohol; and via acyl chloride if known).

6. Classify butyl alcohols: butan-1-ol, butan-2-ol, 2-methylpropan-2-ol as 1°/2°/3° and predict

oxidation products (or none).

7. Write an equation for the addition polymerisation of ethene and identify the repeating unit.

8. Amines: explain why CH_3NH_2 turns damp red litmus blue; write an equation with water or H^+.

9. Two-step synthesis: convert ethene to ethanoic acid (via ethanol). Give reagents/conditions for each step.

10. Two-step synthesis: convert bromoethane to ethanal (via ethanol). Give reagents/conditions for each step.

11. Write the structural formula of the ester from propanoic acid and methanol; name it.

12. State the reagents for dehydration of ethanol and name the organic product.

13. Compare the acidity of ethanol and ethanoic acid (which is stronger acid?) and give one reaction ethanoic acid undergoes that ethanol does not with carbonates.

14. Poly(propene): give the monomer formula and describe the repeating unit.

Further Practice （进阶）· 9

Further Practice

A. Alcohol oxidation and carbonyl tests （醇氧化与羰基检验）

1. State reagents and conditions to obtain propanal from propan-1-ol, and write an equation using [O].

2. State reagents and conditions to obtain propanoic acid from propan-1-ol, and write an equation using [O].

3. Describe how Tollens' reagent distinguishes propanal from propanone, including the observation.

B. Acids, esters and amines （羧酸、酯与胺）

4. Write equations for propanoic acid reacting separately with NaOH and Na_2CO_3.

5. Name and draw the ester formed from methanol and butanoic acid; state the catalyst used.

6. Write equations showing methylamine acting as a Brønsted–Lowry base with water and with hydrochloric acid.

C. Polymerisation and short synthesis （聚合与短合成路线）

7. Draw the repeating unit of poly(propene) and identify the bond in the monomer that opens during polymerisation.

8. Give reagents and conditions for the two-step conversion bromoethane → ethanol → ethanoic acid.

9. Give reagents and conditions for converting ethanol to ethyl ethanoate in two stages by first oxidising part of the ethanol.

Advanced Practice （培优）· 8

Advanced Practice

A. Identification from reaction evidence （反应证据推断结构）

1. Compound A, $C_4H_{10}O$, is oxidised to a carbonyl compound B. B gives no reaction with Tollens' reagent. Deduce a possible structure and name for A and B, and write the oxidation equation.

2. Compound C, C_3H_6O, gives a silver mirror and is oxidised further to D, which effervesces with Na_2CO_3. Identify C and D. Write equations for oxidation of C to D and for D reacting with Na_2CO_3.

B. Synthesis planning and competing routes （路线设计与条件控制）

3. Plan a route from ethene to ethyl ethanoate in no more than three reaction stages. Give all intermediates, reagents and conditions.

4. Starting from propene, devise a route to propanone. State the product and reaction type at each stage.

C. Acyl derivatives and equations （酰基衍生物与方程式）

5. Write equations for ethanoyl chloride reacting separately with ethanol and excess ammonia. Name both organic products.

6. Compare preparation of ethyl ethanoate from ethanoic acid with preparation from ethanoyl chloride in terms of equation, catalyst and reversibility.

D. Polymer structure and monomer deduction （重复单元反推单体）

7. Two candidate monomers are proposed for a polymer whose repeating unit is $-CH_2C(CH_3)_2-$:
   M1: $CH_2{=}C(CH_3)_2$; M2: $CH_3CH{=}CHCH_3$.
   Identify the correct monomer, name it, state the polymerisation type, and explain why the other candidate is inconsistent.

8. Chloroethene forms PVC with repeating unit $-CH_2CHCl-$. Explain why addition polymerisation conserves all monomer atoms, calculate the relative mass of one repeating unit, and find the amount (in moles) of repeating units in 5.30 g of PVC.

14 Analytical Techniques（分析技术）

Infrared spectroscopy (key peaks); mass spectrometry (M^+ simple); elemental analysis and empirical formula.

Rules 规则

- **Infrared (IR)**（红外光谱）: bonds absorb characteristic wavenumbers; used to identify functional groups.
- **Key IR absorptions (AS)**（关键峰）: O–H (acid) broad ∼2500–3000 $\mathrm{cm^{-1}}$; O–H (alcohol) ∼3200–3600 $\mathrm{cm^{-1}}$; C=O ∼1680–1750 $\mathrm{cm^{-1}}$; C–H around 2800–3100 $\mathrm{cm^{-1}}$ (use data book values in exams).
- **Interpreting IR**（识谱）: presence/absence of O–H and C=O distinguishes alcohols, carbonyls, carboxylic acids (acid: broad O–H + C=O).
- **Mass spectrometry**（质谱）: molecular ion peak M^+ gives M_r (approximate molecular mass); fragmentation gives smaller ions.
- **M^+ and isotopes**（分子离子与同位素）: $M+1$, $M+2$ peaks may appear (e.g. ^{13}C; Cl/Br isotope patterns—awareness at AS).
- **Elemental analysis**（元素分析）: percentage composition by mass → moles of atoms → simplest ratio → empirical formula. In combustion analysis of a compound containing C, H and O:

$$\mathrm{C_xH_yO_z} + \left(x + \frac{y}{4} - \frac{z}{2}\right)\mathrm{O_2} \longrightarrow x\,\mathrm{CO_2} + \frac{y}{2}\,\mathrm{H_2O}.$$

  Hence $n(\mathrm{C}) = n(\mathrm{CO_2})$ and $n(\mathrm{H}) = 2n(\mathrm{H_2O})$; oxygen is found by mass difference.
- **Molecular formula from empirical**（由经验式得分子式）: find n such that $n \times M_\mathrm{r}(\text{empirical}) = M_\mathrm{r}$ from M^+ (or given M_r).
- **Combined use**（综合）: IR gives functional groups; MS gives M_r; combustion/elemental data give the empirical formula. A secure identification should satisfy all three sets of evidence, not just one peak.

Worked Examples 例题

Example 1. An IR spectrum shows a strong absorption at about 1700 $\mathrm{cm^{-1}}$ and a very broad band from about 2500–3300 $\mathrm{cm^{-1}}$. Which functional group is suggested?

Carboxylic acid (−COOH): C=O near 1700 $\mathrm{cm^{-1}}$ plus characteristic broad O–H.

Example 2. The mass spectrum of a compound shows a molecular ion at $m/z = 46$. Suggest the M_r and a possible molecular formula for an alcohol.

$M_r = 46$.

Possible: ethanol C_2H_6O ($M_r = 46.0$).

Example 3. A compound contains 52.2% C, 13.0% H and 34.8% O by mass. Find its empirical formula.

Assume 100 g: $n(\mathrm{C}) = \dfrac{52.2}{12.0} = 4.35$, $n(\mathrm{H}) = \dfrac{13.0}{1.0} = 13.0$, $n(\mathrm{O}) = \dfrac{34.8}{16.0} = 2.175$.

Divide by 2.175: ratio $C : H : O = 2 : 6 : 1 \Rightarrow C_2H_6O$.

Basic Drill （入门综合）· 28

1. State what IR spectroscopy is used to identify in organic molecules.

2. Roughly what wavenumber region is associated with C=O stretching?

3. Roughly what wavenumber region is associated with O–H in alcohols?

4. How does the O–H absorption of a carboxylic acid typically appear compared with an alcohol?

5. An IR spectrum has a strong peak near $1720\,\mathrm{cm^{-1}}$ but no broad O–H band. Suggest a functional group class.

6. An IR spectrum has a broad band near $3300\,\mathrm{cm^{-1}}$ and no C=O peak. Suggest a functional group.

7. What does the molecular ion peak M^+ in a mass spectrum tell you?

8. A molecular ion appears at $m/z = 16$. Suggest the identity of the molecule (simple gas).

9. A molecular ion appears at $m/z = 28$. Suggest two possible molecules (N_2 or CO).

10. Define empirical formula.

11. A compound is 85.7% C and 14.3% H by mass. Find the empirical formula.

12. Empirical formula CH_2; $M_r = 56$. Find the molecular formula.

13. Empirical formula CH_2O; $M_r = 60$. Find the molecular formula.

14. Elemental analysis: 40.0% C, 6.7% H, 53.3% O. Find the empirical formula.

15. Why might a mass spectrum show a small peak at $M + 1$?

16. State one reason IR alone may not uniquely identify a compound.

17. Which technique gives M_r more directly: IR or mass spectrometry?

18. A carboxylic acid and an alcohol both may show O–H in IR. What extra IR feature points to the acid?

19. Propane has $M_r = 44$. Where would you expect M^+?

20. Ethanol has $M_\mathrm{r} = 46$. Where would you expect M^+?

21. Fragmentation: what does a peak at m/z lower than M^+ often represent?

22. A hydrocarbon empirical formula is CH. $M_\mathrm{r} = 78$. Find the molecular formula.

23. Calculate percentage by mass of C in ethene $\mathrm{C_2H_4}$.

24. From combustion analysis, oxides of C and H are collected. Which elements' amounts are thus found?

25. Match: IR $\leftrightarrow$ functional groups; MS $\leftrightarrow$ …

26. A ketone shows strong IR absorption near $1715\,\mathrm{cm^{-1}}$. Which bond causes this?

27. Would you expect a strong broad IR O–H band for dry ethoxyethane (an ether)? Explain.

28. Compound X has empirical formula $\mathrm{C_2H_4O}$ and $M^+ = 88$. Find the molecular formula.

Practice （综合练习）· 12

Practice

1. Distinguish propan-1-ol, propanal and propanoic acid using IR absorptions you would expect for each.

2. A compound contains 54.5% C, 9.1% H and 36.4% O. $M_{\mathrm{r}} = 88$. Find empirical and molecular formulas.

3. The mass spectrum of chloroethane shows peaks near M and $M + 2$. Explain briefly in terms of chlorine isotopes.

4. A molecular ion is observed at $m/z = 58$. Suggest two possible oxygen-containing formulas and how IR could distinguish them (ketone vs other).

5. Elemental analysis of an amine gives 38.7% C, 16.1% H, 45.2% N. Find the empirical formula.

6. Explain how M^+ and elemental analysis together give a molecular formula.

7. An IR spectrum shows C=O and a sharp (not very broad acid) O–H. Could this be a carboxylic acid? What else might it be?

8. Calculate the empirical formula of a compound with composition 92.3% C, 7.7% H. If $M_{\mathrm{r}} = 78$, find the molecular formula and suggest a name.

9. Why is the molecular ion peak sometimes very weak or absent? (simple: unstable molecular ion / easy fragmentation)

10. A student obtains IR of ethanoic acid. Identify two characteristic absorptions and assign them.

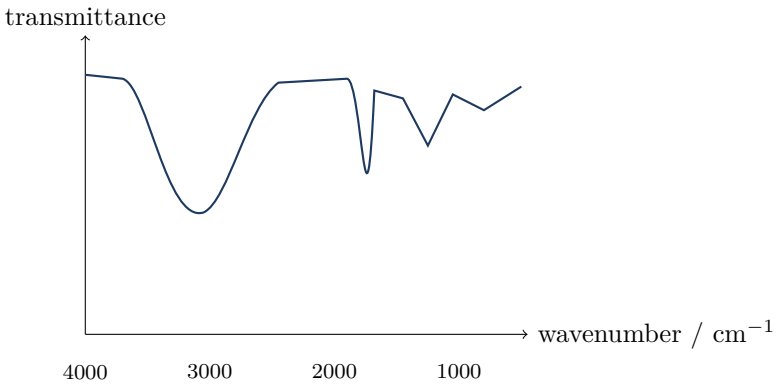
11. From percentages $C = 40.0\%$, $H = 6.7\%$, $O = 53.3\%$ and $M^+ = 180$, find molecular formula.

12. Outline how you would use IR and MS to confirm a sample is ethanol rather than ethanal.

Further Practice （进阶）· 9

Further Practice

A. Reading an IR spectrum （红外谱图判读）



1. The schematic IR spectrum has a very broad absorption in the 2500–3300 cm^{-1} region and a strong absorption near 1710 cm^{-1}. Identify both bonds and the functional group indicated.

2. State how the spectrum of an alcohol would differ from the spectrum shown.

3. If only the strong absorption near 1710 cm^{-1} were considered, explain why propanal and propanone could not be distinguished completely.

B. Molecular ions and isotope patterns （分子离子峰与同位素峰）

4. A molecular ion peak occurs at $m/z = 60$. State the relative molecular mass and suggest one possible oxygen-containing molecular formula.

5. Explain why chloroethane gives molecular-ion peaks at M and $M+2$ in an approximate $3:1$ ratio.

6. A bromine-containing molecular ion gives peaks at $m/z = 108$ and 110 of similar height. Explain the pattern.

C. Empirical and molecular formulae （元素分析与分子式）

7. A compound contains 54.5% C, 9.1% H and 36.4% O. Determine its empirical formula.

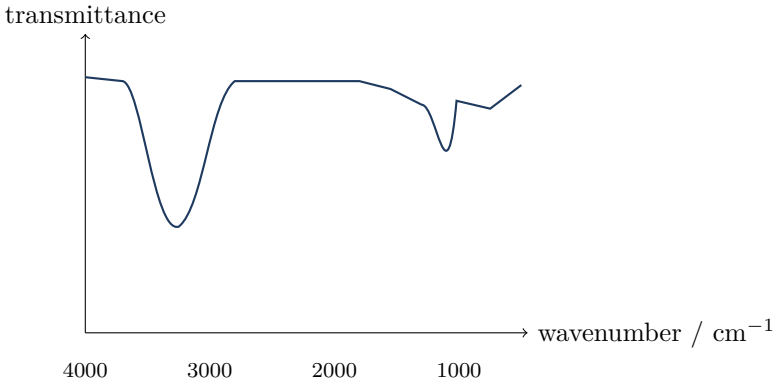
8. The compound in Q7 has $M^+ = 88$. Determine its molecular formula.

9. A hydrocarbon contains 85.7% C by mass and has a molecular ion at $m/z = 84$. Determine its empirical and molecular formulae.

Advanced Practice （培优）· 8

Advanced Practice

A. Combined spectral identification （红外与质谱联合推断）



1. Compound X has the IR spectrum shown, $M^+ = 60$, and composition 60.0% C, 13.3% H and 26.7% O. Determine its empirical formula, molecular formula, functional group and one possible identity.

2. Another compound Y also has $M^+ = 60$ but shows a strong absorption near $1715\,\mathrm{cm^{-1}}$ and no broad O–H absorption. Suggest a structure for Y and explain why Y is not an isomer of X.

B. Quantitative elemental analysis （定量元素分析）

3. Complete combustion of 0.880 g of a compound containing only C, H and O produces 1.760 g CO_2 and 0.720 g H_2O. Determine the empirical formula.

4. The compound in Q3 has a molecular ion at $m/z = 88$ and a strong C=O absorption. Determine its molecular formula and suggest two possible functional-group classes.

C. Isotope and fragment constraints （同位素与碎片峰约束）

5. A compound gives molecular-ion peaks at $m/z = 78$ and 80 in a $3:1$ ratio. Deduce which halogen is present and suggest a possible molecular formula.

6. Explain why a strong fragment peak cannot by itself be used as the relative molecular mass, and state how the molecular ion is recognised.

D. Evaluating identification evidence （鉴定证据的充分性）

7. Two compounds have the same empirical formula, molecular ion and C=O absorption. Explain why they may still be different compounds and name one additional chemical test that could distinguish an aldehyde from a ketone.

8. Design a minimal evidence set using elemental analysis, mass spectrometry and IR to confirm that an unknown is ethanoic acid rather than another compound with $M_\mathrm{r} = 60$.