Week 32

A-Level Chemistry

Homework bundle for this week

本周作业合集

exercises.lol

Contents 目录

A-Level Chemistry Name: _____ Score: _____

1. The lattice energy is:
   - ☐ the (negative) energy change when one mole of solid lattice forms from gaseous ions
   - ☐ always positive
   - ☐ the energy to make gaseous atoms
   - ☐ the energy to remove an electron
2. The enthalpy change of hydration is:
   - ☐ exothermic — energy released when gaseous ions are surrounded by water
   - ☐ always endothermic
   - ☐ the energy to break the lattice
   - ☐ zero
3. Entropy is a measure of:
   - ☐ the number of ways the particles and their energy can be arranged (disorder)
   - ☐ the energy released by a reaction
   - ☐ the temperature
   - ☐ the mass of a substance
4. The Gibbs free energy change is calculated as:
   - ☐ $\Delta G = \Delta H - T\Delta S$
   - ☐ $\Delta G = \Delta H + T\Delta S$
   - ☐ $\Delta G = T\Delta S - \Delta H$
   - ☐ $\Delta G = \Delta H \times \Delta S$
5. Going down Group 17, the first electron affinity becomes:
   - ☐ less exothermic, because the larger atom attracts the electron less strongly
   - ☐ more exothermic
   - ☐ always positive
   - ☐ unchanged
6. The enthalpy change when one mole of gaseous ions dissolves in water is the enthalpy of _____.

_____

7. The enthalpy change when one mole of an ionic solid forms from gaseous ions is the _____ energy.

   _____

8. Enthalpy of solution can be found from lattice energy and hydration enthalpies using Hess's law.

   ☐ True ☐ False

9. Entropy increases when a solid melts or a gas is formed.

   ☐ True ☐ False

10. A reaction is feasible when the Gibbs free energy change ΔG is _____.

    _____

23. Chemical energetics

Answers · 答案

A-Level Chemistry

1. **the (negative) energy change when one mole of solid lattice forms from gaseous ions**

   Strong ionic bonds form, so lattice energy is exothermic (negative).

2. **exothermic — energy released when gaseous ions are surrounded by water**

   Water molecules are attracted to the ions, releasing energy, so hydration is exothermic.

3. **the number of ways the particles and their energy can be arranged (disorder)**

   More ways to arrange particles/energy means higher entropy (more disorder).

4. **$\Delta G = \Delta H - T\Delta S$**

   $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$, with T in kelvin.

5. **less exothermic, because the larger atom attracts the electron less strongly**

   A bigger atom pulls the incoming electron in less strongly, so EA is less exothermic.

6. **hydration**

   More negative for smaller, more highly charged ions.

7. **lattice**

   Born–Haber cycles use Hess's law to find it.

8. **True**

   A Hess cycle links them.

9. **True**

   More disorder means higher entropy.

10. **negative**

    $\Delta G = \Delta H - T\Delta S$.

A-LEVEL CHEMISTRY • EXERCISE SHEET

23.1 Lattice energy and Born-Haber cycles

Name: _____ Class: _____ Date: _____ **Total: 11 marks**

KEY WORDS enthalpy change 焓变 • lattice energy 晶格能 • electron affinity 电子亲和能 • atomisation 原子化 • enthalpy change of atomisation 原子化焓变

Objective

Build the skills to answer exam questions on **lattice energy and Born–Haber cycles** — **atomisation** 原子化, **lattice energy** 晶格能, **electron affinity** 电子亲和能, and Born–Haber calculations.

You must be able to:

- define enthalpy of atomisation, lattice energy and first electron affinity
- explain the trends in electron affinity down Groups 16 and 17
- construct and use Born–Haber cycles to calculate an unknown step

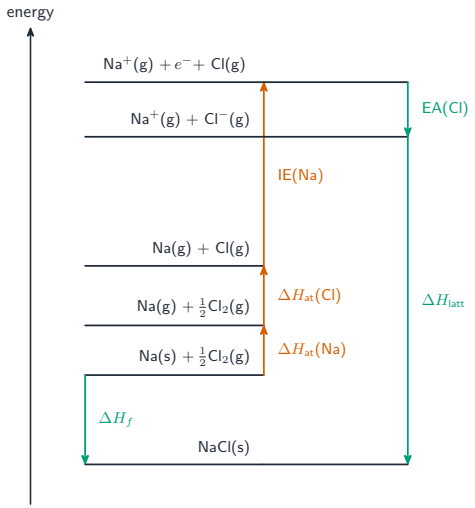
1 Worked examples

The key terms

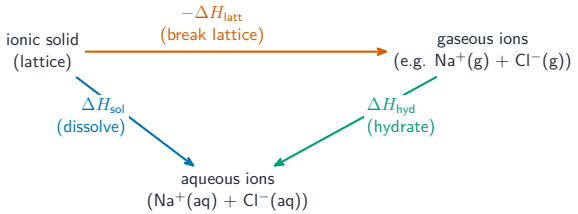
- $\Delta H_{\mathbf{at}}$ — energy to make 1 mol of gaseous atoms from an element (always +).
- $\Delta H_{\mathbf{latt}}$ — change when 1 mol of ionic solid forms from gaseous ions (always −).
- **first electron affinity (EA)** — change when 1 mol of gaseous atoms each gain an electron → 1 mol of 1− ions (usually −).

EA becomes **less exothermic** down Groups 16/17 (larger atom attracts the electron less).

Born–Haber cycle



By Hess's law, the direct route (ΔHf) equals the sum of the steps round the cycle — solve for any unknown



$$\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$$

Born-Haber alongside the dissolving cycle

For NaCl: $\Delta H_f = \Delta H_{\text{at}}(\text{Na}) + \text{IE}_1(\text{Na}) + \Delta H_{\text{at}}(\text{Cl}) + \text{EA}_1(\text{Cl}) + \Delta H_{\text{latt}}$.

2 Practice

2.1 Define **lattice energy**. [2]

2.2 State why the first electron affinity of chlorine is exothermic. [1]

3 Exam-style questions

3.1 The enthalpy change of atomisation is always: [1]

- **A** negative
- **B** positive
- **C** zero
- **D** equal to the lattice energy

3.2 Down Group 17, the first electron affinity becomes: [1]

- **A** more exothermic
- **B** less exothermic
- **C** zero
- **D** positive

3.3 Use these data (kJ mol^{-1}) for NaCl: $\Delta H_f = -411$; $\Delta H_{at}(Na) = +107$; $IE_1(Na) = +496$; $\Delta H_{at}(Cl) = +122$; $EA_1(Cl) = -349$.

(a) Calculate the lattice energy of NaCl. [3]

(b) State why the lattice energy is negative. [1]

3.4 Explain why the lattice energy of MgO is more exothermic than that of NaCl. [2]

Solutions

2.1 the enthalpy change when 1 mol of an ionic solid (lattice) forms from its gaseous ions.

2.2 the atom gains an electron, and the attraction of the nucleus for the added electron releases energy.

3.1 B. Atomisation always requires energy (bonds break), so it is positive.

3.2 B. EA becomes less exothermic down the group.

3.3 (a) $\Delta H_{\text{latt}} = \Delta H_f - [\Delta H_{\text{at}}(\text{Na}) + \text{IE}_1 + \Delta H_{\text{at}}(\text{Cl}) + \text{EA}_1]$; $= -411 - [107 + 496 + 122 - 349]$ $= -411 - 376 = -787$ kJ mol^{-1}.
(b) strong electrostatic attractions form between the oppositely charged ions (energy released).

3.4 MgO has ions with higher charges (2+/2−) and smaller radii than NaCl; so the electrostatic attraction is much stronger, making the lattice energy more exothermic.

4 **(a)** Define enthalpy change of atomisation, ΔH_{at}.

.....

..... [1]

(b) Define first electron affinity, EA.

.....

..... [1]

(c) Explain why the first electron affinity of chlorine is more exothermic than the first electron affinity of iodine.

.....

.....

.....

..... [2]

(d) The enthalpy change for the reaction $Cl_2(g) + 2e^- \rightarrow 2Cl^-(g)$ is $-486\,kJ\,mol^{-1}$.

The first electron affinity of chlorine is $-364\,kJ\,mol^{-1}$.

Calculate the enthalpy change of atomisation of chlorine.

ΔH_{at} of chlorine = ... kJ mol^{-1} [2]

[Total: 6]

2 **(a)** The Group 2 sulfates and the Group 2 chromates show similar trends in solubility.

Suggest the trend in the solubility of the Group 2 chromates down the group.

Explain your answer.

.....

.....

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.....

.....

..... [4]

(b) Silver(I) chromate, Ag_2CrO_4, is sparingly soluble in water.

(i) Write an ionic equation to show the equilibrium between solid Ag_2CrO_4 and its aqueous solution.

Include state symbols.

..... [1]

(ii) The value of the solubility product, K_{sp}, of Ag_2CrO_4 is 1.12×10^{-12} at 298 K.

Calculate the equilibrium concentration of Ag^+, in mol dm^{-3}, in a saturated solution of Ag_2CrO_4 at 298 K.

equilibrium concentration of Ag^+ = ... mol dm^{-3} [3]

(c) The hydrogenchromate ion, $HCrO_4^-$, is a weak acid. The pK_a of $HCrO_4^-$ is 6.49.

(i) Calculate the pH of a 0.0250 mol dm^{-3} $HCrO_4^-$ solution.

pH = ... [2]

(ii) $HCrO_4^-$ can show amphoteric behaviour.

State the formula of:

- the conjugate acid of $HCrO_4^-$ .....
- the conjugate base of $HCrO_4^-$. .....

[1]

(d) Table 2.1 shows some energy changes.

Table 2.1

energy change	value / kJ mol^{-1}
first ionisation energy of silver	+731
second ionisation energy of silver	+2074
first ionisation energy of sulfur	+1000
second ionisation energy of sulfur	+2251
first electron affinity of sulfur	–200
second electron affinity of sulfur	+532
enthalpy change of atomisation of sulfur	+279
enthalpy change of formation of silver(I) sulfide, $Ag_2S(s)$	–33
lattice energy of silver(I) sulfide, $Ag_2S(s)$	–2677

(i) Define the term first electron affinity.

.....

.....

..... [1]

(ii) Explain why the value for the second electron affinity of sulfur is positive.

.....

.....

..... [1]

(iii) Construct an equation for the lattice energy of Ag_2S.

Include state symbols.

..... [1]

(iv) Calculate the enthalpy change of atomisation, ΔH_{at}, in kJ mol^{-1}, of silver using relevant data from Table 2.1.

It may be helpful to draw a labelled Born–Haber cycle.

Show your working.

ΔH_{at} of silver = ... kJ mol^{-1} [3]

(e) Suggest how the magnitude for the lattice energy of $Ag_2S(s)$ differs from the lattice energy of $Cu_2S(s)$.

Explain your answer.

.....

.....

..... [1]

[Total: 18]

23.2 Enthalpies of solution and hydration

Name: _____ Class: _____ Date: _____ **Total: 23 marks**

KEY WORDS enthalpy change of hydration 水合焓变 • enthalpy change of solution 溶解焓变 • enthalpy change 焓变 • lattice energy 晶格能

Objective

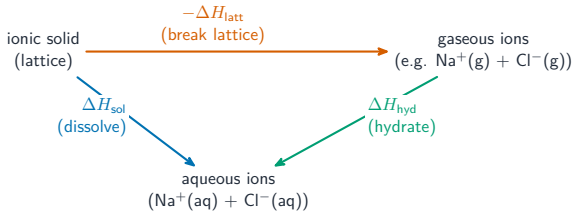
Build the skills to answer exam questions on **enthalpies of solution and hydration** — $\Delta H_{\mathbf{sol}}$ 溶解焓 and $\Delta H_{\mathbf{hyd}}$ 水合焓, the energy cycle, and the effect of ionic charge and radius.

You must be able to:

- define enthalpy of solution and enthalpy of hydration
- construct and use an energy cycle linking them with lattice energy
- explain the effect of ionic charge and radius on the magnitude of hydration enthalpy

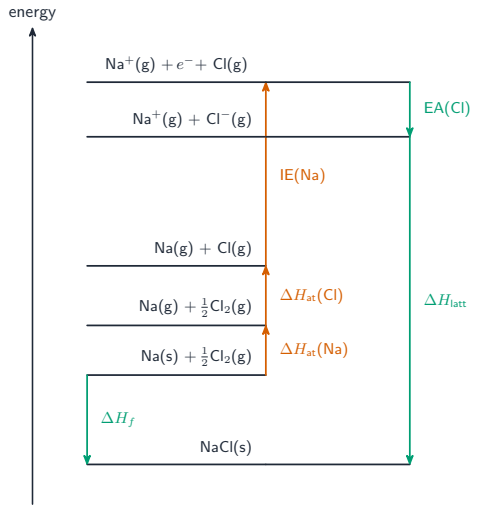
1 Worked examples

The dissolving cycle



$$\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$$

Dissolving = breaking the lattice ($-\Delta H_{latt}$) then hydrating the ions (ΔH_{hyd})



Lattice energy from a Born-Haber cycle

$$\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \sum \Delta H_{\text{hyd}}$$

- $\Delta H_{\mathbf{hyd}}$ — change when 1 mol of gaseous ions is dissolved in water (always −, water attracts the ion).
- $\Delta H_{\mathbf{sol}}$ — change when 1 mol of solute dissolves to form a solution.

Charge and radius

Hydration is **more exothermic** for ions with a **higher charge** and a **smaller radius** (higher charge density attracts the polar water more strongly).

2 Practice

2.1 Define **enthalpy change of hydration**. [2]

2.2 State **two** factors that make hydration of an ion more exothermic. [2]

3 Exam-style questions

3.1 Enthalpy of hydration is most exothermic for: [1]

- **A** a large, singly-charged ion
- **B** a small, highly-charged ion
- **C** a large, highly-charged ion
- **D** a neutral atom

3.2 Which equation correctly links the enthalpies? [1]

- **A** $\Delta H_{\text{sol}} = \Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$
- **B** $\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$
- **C** $\Delta H_{\text{sol}} = \Delta H_{\text{latt}} - \Delta H_{\text{hyd}}$
- **D** $\Delta H_{\text{sol}} = -\Delta H_{\text{hyd}}$

3.3 For a salt, $\Delta H_{\text{latt}} = -780$ kJ mol^{-1} and the sum of the hydration enthalpies of its ions is -820 kJ mol^{-1}.

(a) Calculate the enthalpy change of solution. [2]

(b) State whether the salt dissolving is exothermic or endothermic. [1]

3.4 Explain why the hydration enthalpy of Mg^{2+} is more exothermic than that of Na^{+}. [2]

3.5 The table gives data for two Group 2 chlorides.

	$\Delta H^{\ominus}_{\text{LE}}$ / kJ mol^{-1}	$\Delta H^{\ominus}_{\text{sol}}$ / kJ mol^{-1}
$MgCl_2$	-2526	-155
$CaCl_2$	-2258	-83

$\Delta H^{\ominus}_{\text{hyd}}(Cl^-) = -364$ kJ mol^{-1}.

(a) Draw the energy cycle linking lattice enthalpy, enthalpy of solution and the enthalpies of hydration, and use it to calculate $\Delta H^{\ominus}_{\text{hyd}}(Mg^{2+})$. [5]

(b) Explain why $\Delta H^{\ominus}_{\text{hyd}}(Mg^{2+})$ is more exothermic than $\Delta H^{\ominus}_{\text{hyd}}(Ca^{2+})$. [3]

(c) Both dissolvings are exothermic, yet the magnitudes are small compared with the individual terms. Explain why. [2]

(d) State one reason why a calculated lattice enthalpy from a purely ionic model may differ from the experimental value. [2]

Solutions

2.1 the enthalpy change when 1 mol of gaseous ions is dissolved in (a large amount of) water.

2.2 a higher ionic charge; a smaller ionic radius.

3.1 B. A small, highly-charged ion has the highest charge density, so the most exothermic hydration.

3.2 B. $\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$.

3.3 (a) $\Delta H_{\text{sol}} = -(-780) + (-820) = +780 - 820 = -40$ kJ mol^{-1}.
(b) exothermic.

3.4 Mg^{2+} has a higher charge and a smaller radius than Na^{+}, so a higher charge density that attracts the polar water molecules more strongly — more exothermic hydration.

3.5 (a) The cycle goes: solid $\rightarrow$ gaseous ions ($-\Delta H_{\text{LE}}$) $\rightarrow$ aqueous ions ($\sum \Delta H_{\text{hyd}}$), and directly solid $\rightarrow$ aqueous ions (ΔH_{sol}). So $\Delta H_{\text{sol}} = -\Delta H_{\text{LE}} + \Delta H_{\text{hyd}}(Mg^{2+}) + 2\Delta H_{\text{hyd}}(Cl^{-})$; $-155 = 2526 + \Delta H_{\text{hyd}}(Mg^{2+}) + 2(-364)$; $\Delta H_{\text{hyd}}(Mg^{2+}) = -155 - 2526 + 728 = -1953$ kJ mol^{-1}.
(b) Mg^{2+} is the smaller ion but carries the same 2+ charge, so it has the greater charge density; it attracts the $\delta-$ oxygen of the water molecules more strongly, releasing more energy per mole.
(c) ΔH_{sol} is the small **difference** between two very large opposing terms — the energy needed to pull the lattice apart and the energy released on hydration; they nearly cancel, so the net value is small.
(d) The purely ionic model assumes perfect spheres with no covalent character; a small, highly charged cation polarises a large anion, giving some covalent character and a larger experimental lattice enthalpy than calculated.

1 **(a)** Solutions of Group 2 hydrogencarbonates, $M(HCO_3)_2$, decompose on heating to give the corresponding metal carbonate, carbon dioxide and water.

(i) Write an equation for the decomposition of strontium hydrogencarbonate, $Sr(HCO_3)_2$.

..... [1]

(ii) The thermal stability of Group 2 carbonates increases down the group.

Explain this trend.

.....

.....

.....

..... [2]

(b) The hydroxides and fluorides of Group 2 elements show similar trends in solubility.

Describe the trend in the solubility of the fluorides of calcium, strontium and barium.

Explain your answer.

....
least soluble most soluble

explanation .....

.....

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..... [4]

(c) **(i)** Define enthalpy change of hydration, ΔH_{hyd}.

.....

.....

..... [1]

(ii) State the main factors that affect the magnitude of enthalpy change of hydration.

Explain your answer.

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.....

..... [2]

(d) Table 1.1 shows various energy changes.

Table 1.1

energy change	value / kJ mol^{-1}
lattice energy of MgF_2	–2957
enthalpy change of hydration, ΔH_{hyd}, of Mg^{2+}	–1926
enthalpy change of hydration, ΔH_{hyd}, of F^-	–505

Use data from Table 1.1 to calculate the enthalpy change of solution, ΔH_{sol}, for $MgF_2(s)$.

It may be helpful to draw a labelled energy cycle. Show your working.

ΔH_{sol} of $MgF_2(s)$ = ... kJ mol^{-1} [2]

(e) Mercury(I) fluoride, Hg_2F_2, is sparingly soluble in water.

The cation in Hg_2F_2 exists as the diatomic ion Hg_2^{2+} with a covalent Hg–Hg bond.

(i) Write the expression for the solubility product, K_{sp}, of Hg_2F_2. Include the units.

K_{sp} =

units .....
[2]

(ii) The solubility of Hg_2F_2 is 9.20×10^{-3} mol dm^{-3} at 298 K.

Calculate the value of K_{sp} of Hg_2F_2 at 298 K.

K_{sp} = ... [1]

[Total: 15]

1 Magnesium nitrate, $Mg(NO_3)_2$, and strontium nitrate, $Sr(NO_3)_2$, both decompose when heated to form the metal oxide and a mixture of gases.

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

..... [1]

(b) State which of $Mg(NO_3)_2$ or $Sr(NO_3)_2$ decomposes at a lower temperature.

Explain your answer.

compound that decomposes at a lower temperature .....

explanation .....

.....

.....

.....

.....

[2]

(c) Magnesium oxide, MgO, and strontium oxide, SrO, both react with dilute sulfuric acid.

MgO forms a soluble salt, **A**.

SrO forms an insoluble salt, **B**.

(i) Identify the products formed when MgO reacts with dilute sulfuric acid.

..... [1]

(ii) Explain why **A** is more soluble than **B**.

.....

.....

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.....

.....

..... [3]

[Total: 7]

2 **(a)** The Group 2 sulfates and the Group 2 chromates show similar trends in solubility.

Suggest the trend in the solubility of the Group 2 chromates down the group.

Explain your answer.

.....

.....

.....

.....

.....

..... [4]

(b) Silver(I) chromate, Ag_2CrO_4, is sparingly soluble in water.

(i) Write an ionic equation to show the equilibrium between solid Ag_2CrO_4 and its aqueous solution.

Include state symbols.

..... [1]

(ii) The value of the solubility product, K_{sp}, of Ag_2CrO_4 is 1.12×10^{-12} at 298 K.

Calculate the equilibrium concentration of Ag^+, in $mol\,dm^{-3}$, in a saturated solution of Ag_2CrO_4 at 298 K.

equilibrium concentration of Ag^+ = ... $mol\,dm^{-3}$ [3]

(c) The hydrogenchromate ion, $HCrO_4^-$, is a weak acid. The pK_a of $HCrO_4^-$ is 6.49.

(i) Calculate the pH of a 0.0250 mol dm^{-3} $HCrO_4^-$ solution.

pH = ... [2]

(ii) $HCrO_4^-$ can show amphoteric behaviour.

State the formula of:

- the conjugate acid of $HCrO_4^-$ .....
- the conjugate base of $HCrO_4^-$. .....

[1]

(d) Table 2.1 shows some energy changes.

Table 2.1

energy change	value / kJ mol^{-1}
first ionisation energy of silver	+731
second ionisation energy of silver	+2074
first ionisation energy of sulfur	+1000
second ionisation energy of sulfur	+2251
first electron affinity of sulfur	–200
second electron affinity of sulfur	+532
enthalpy change of atomisation of sulfur	+279
enthalpy change of formation of silver(I) sulfide, $Ag_2S(s)$	–33
lattice energy of silver(I) sulfide, $Ag_2S(s)$	–2677

(i) Define the term first electron affinity.

.....

.....

..... [1]

(ii) Explain why the value for the second electron affinity of sulfur is positive.

.....

.....

..... [1]

(iii) Construct an equation for the lattice energy of Ag_2S.

Include state symbols.

..... [1]

(iv) Calculate the enthalpy change of atomisation, ΔH_{at}, in kJ mol^{-1}, of silver using relevant data from Table 2.1.

It may be helpful to draw a labelled Born–Haber cycle.

Show your working.

ΔH_{at} of silver = ... kJ mol^{-1} [3]

(e) Suggest how the magnitude for the lattice energy of $Ag_2S(s)$ differs from the lattice energy of $Cu_2S(s)$.

Explain your answer.

.....

.....

..... [1]

[Total: 18]

A-LEVEL CHEMISTRY • EXERCISE SHEET

23.3 Entropy change

Name: _____ Class: _____ Date: _____ **Total: 10 marks**

KEY WORDS entropy 熵 • entropy change 熵变 • feasibility 可行性

Objective

Build the skills to answer exam questions on **entropy change** 熵变 — what **entropy** 熵 (S) is, predicting the sign of ΔS, and calculating it.

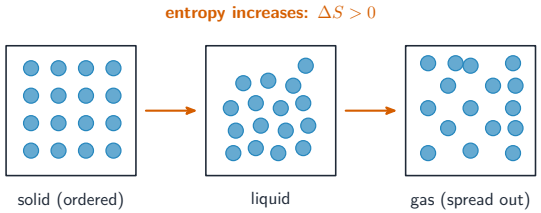
You must be able to:

- define entropy as the number of ways the particles and energy can be arranged
- predict and explain the sign of ΔS for state changes and gas-number changes
- calculate ΔS from standard entropies

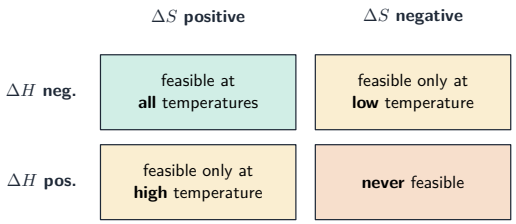
1 Worked examples

Entropy and its sign

Entropy S = the number of possible arrangements of the particles and their energy. More disorder = higher entropy.



Entropy rises on melting, boiling and dissolving, and when more gas molecules are made ($+\Delta S$)



$\Delta G = \Delta H - T\Delta S$; **feasible when** $\Delta G \leq 0$

Entropy feeds the Gibbs feasibility test

- **+ΔS:** solid → liquid → gas; dissolving; more gas molecules formed; higher temperature.
- **−ΔS:** the reverse (fewer gas molecules, freezing).

■ **Calculating ΔS**

$$\Delta S^{\ominus} = \sum S^{\ominus}(\text{products}) - \sum S^{\ominus}(\text{reactants})$$

(units J $K^{-1}mol^{-1}$).

2 Practice

2.1 Define **entropy**. [1]

2.2 State the sign of ΔS when a gas condenses to a liquid. [1]

3 Exam-style questions

3.1 Which change has the **largest positive** entropy change? [1]

- **A** freezing of water
- **B** $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
- **C** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
- **D** condensing steam

3.2 The entropy of a substance is **highest** when it is a: [1]

- **A** solid
- **B** liquid
- **C** gas
- **D** crystal

3.3 For $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, the standard entropies (J $K^{-1}mol^{-1}$) are: N_2 192, H_2 131, NH_3 193.

(a) Calculate $\Delta S^{\ominus}$ for the reaction. [3]

(b) Explain the sign of your answer. [1]

3.4 Explain why dissolving a solid in water usually has a positive entropy change. [2]

Solutions

2.1 the number of possible arrangements of the particles and their energy in a system.

2.2 negative.

3.1 B. A solid producing a gas gives the largest increase in entropy.

3.2 C. A gas has the highest entropy.

3.3 (a) $\Delta S^{\ominus} = \sum S(\text{products}) - \sum S(\text{reactants})$; $= 2(193) - [192 + 3(131)] = 386 - 585 = -199$ J $K^{-1}mol^{-1}$.
(b) negative — 4 moles of gas become 2 moles, so the disorder decreases.

3.4 the solid lattice breaks up and its ions/molecules spread out and mix among the water molecules, giving many more arrangements, so the entropy increases.

3 **(a)** Define the term entropy.

.....

..... [1]

(b) **(i)** Place **one** tick (✓) in each row of Table 3.1 to show the sign of the entropy change, ΔS, for each process.

Table 3.1

process	ΔS is negative	ΔS is positive
steam condensing into water		
solid KC*l* dissolving in water		

[1]

(ii) Chlorine trifluoride, ClF_3, decomposes on heating into its elements, as shown.

reaction 1 $2ClF_3(g) \rightarrow Cl_2(g) + 3F_2(g)$

Standard entropies are shown in Table 3.2.

Table 3.2

substance	$ClF_3(g)$	$Cl_2(g)$	$F_2(g)$
$S^{\ominus}$ / J K^{-1} mol^{-1}	+281.6	+223.1	+203.0

Calculate the standard entropy change, $\Delta S^{\ominus}$, in J K^{-1} mol^{-1}, for reaction 1.

$\Delta S^{\ominus}$ for reaction 1 = ... J K^{-1} mol^{-1} [2]

(c) Group 2 carbonates decompose on heating. The decomposition for one of the Group 2 carbonates, MCO_3, is shown in reaction 2.

reaction 2 $MCO_3(s) \rightarrow MO(s) + CO_2(g)$

(i) Predict the sign of the entropy change, ΔS, for reaction 2.

Explain your answer.

.....

..... [1]

(ii) The Gibbs equation is shown.

$$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$$

Fig. 3.1 shows values of the Gibbs free energy change, $\Delta G^{\ominus}$, in kJ mol^{-1}, at different temperatures, T, in K, for reaction 2.

Assume $\Delta H^{\ominus}$ and $\Delta S^{\ominus}$ values for this reaction remain constant over this temperature range.

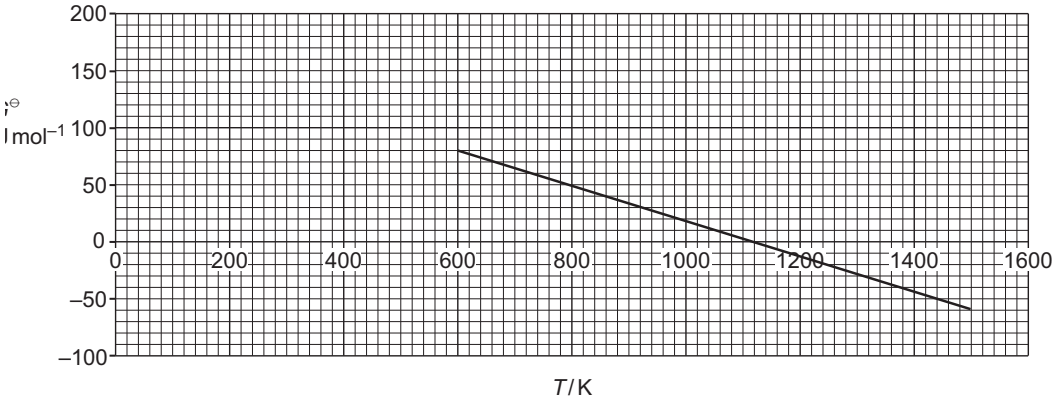


Fig. 3.1

Use the gradient and intercept on the y-axis in Fig. 3.1 and the Gibbs equation to determine:

- $\Delta S^{\ominus}$, in J K^{-1} mol^{-1}, for reaction 2
- the minimum temperature, T, in K, at which the reaction is feasible
- $\Delta H^{\ominus}$, in kJ mol^{-1}, for reaction 2.

$\Delta S^{\ominus}$ for reaction 2 = ... J K^{-1} mol^{-1}

minimum temperature, T = ... K

$\Delta H^{\ominus}$ for r ɔn 2 = ... kJ mol^{-1}
[4]

[Total: 9]

A-LEVEL CHEMISTRY • EXERCISE SHEET

23.4 Gibbs free energy change

Name: _____ Class: _____ Date: _____ **Total: 11 marks**

KEY WORDS gibbs free energy change 吉布斯自由能变 • gibbs free energy 吉布斯自由能 • feasibility 可行性 • entropy 熵 • entropy change 熵变

Objective

Build the skills to answer exam questions on **Gibbs free energy change** 吉布斯自由能变 — using $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$ to judge **feasibility** 可行性.

You must be able to:

- state and use the Gibbs equation
- use the sign of ΔG to decide whether a reaction is feasible
- predict the effect of temperature on feasibility

1 Worked examples

■ **The Gibbs equation**

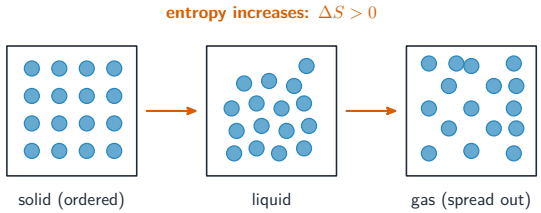
$$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$$

A reaction is **feasible** when $\Delta G \leq 0$.

	ΔS **positive**	ΔS **negative**
ΔH **neg.**	feasible at **all** temperatures	feasible only at **low** temperature
ΔH **pos.**	feasible only at **high** temperature	**never** feasible

$\Delta G = \Delta H - T\Delta S$; **feasible when** $\Delta G \leq 0$

Feasibility depends on temperature: a $+\Delta S$ reaction becomes feasible above a certain T (where $T\Delta S$ outweighs ΔH)



Gibbs combines enthalpy and entropy changes

Watch the units: ΔH is in kJ mol^{-1} but ΔS is in J K^{-1} mol^{-1} — divide ΔS by 1000 (or multiply ΔH by 1000) before subtracting.

Example. ΔH = +178 kJ mol^{-1}, ΔS = +160 J K^{-1} mol^{-1}. At 298 K: $\Delta G = 178 - 298(0.160) = 178 - 47.7 = +130$ kJ mol^{-1} (not feasible). It becomes feasible above $T = \Delta H/\Delta S = 178/0.160 = 1113$ K.

2 Practice

2.1 State the Gibbs equation. [1]

2.2 State the condition on ΔG for a reaction to be feasible. [1]

3 Exam-style questions

3.1 A reaction has ΔH negative and ΔS positive. It is: [1]

- **A** feasible at all temperatures
- **B** feasible only at high temperature
- **C** feasible only at low temperature
- **D** never feasible

3.2 In the Gibbs equation, the temperature T must be in: [1]

- **A** °C
- **B** kelvin
- **C** joules
- **D** kJ

3.3 A reaction has $\Delta H = +120$ kJ mol^{-1} and $\Delta S = +200$ J K^{-1}mol^{-1}.

(a) Calculate ΔG at 298 K and state whether the reaction is feasible. [3]

(b) Calculate the minimum temperature at which the reaction becomes feasible. [2]

3.4 Explain why an endothermic reaction with a positive entropy change becomes feasible only above a certain temperature. [2]

Solutions

2.1 $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$.

2.2 $\Delta G \leq 0$ (negative or zero).

3.1 A. With $-\Delta$H and $+\Delta$S, ΔG is negative at all temperatures.

3.2 B. Temperature must be in kelvin.

3.3 (a) $\Delta G = 120 - 298(0.200) = 120 - 59.6 = +60.4$ kJ mol^{-1}; not feasible (positive ΔG).
(b) feasible when $\Delta G = 0$: $T = \Delta H/\Delta S = 120/0.200 = 600$ K.

3.4 at low T the $T\Delta S$ term is small, so $\Delta G = \Delta H - T\Delta S$ stays positive; raising T increases $T\Delta S$ until it exceeds ΔH, making ΔG negative and the reaction feasible.

3 **(a)** Define the term entropy.

.....

..... [1]

(b) **(i)** Place **one** tick (✓) in each row of Table 3.1 to show the sign of the entropy change, ΔS, for each process.

Table 3.1

process	ΔS is negative	ΔS is positive
steam condensing into water		
solid KC*l* dissolving in water		

[1]

(ii) Chlorine trifluoride, C*l*F$_3$, decomposes on heating into its elements, as shown.

reaction 1 $2ClF_3(g) \rightarrow Cl_2(g) + 3F_2(g)$

Standard entropies are shown in Table 3.2.

Table 3.2

substance	C*l*F$_3$(g)	C*l*$_2$(g)	F$_2$(g)
$S^{\ominus}$ / J K^{-1} mol^{-1}	+281.6	+223.1	+203.0

Calculate the standard entropy change, $\Delta S^{\ominus}$, in J K^{-1} mol^{-1}, for reaction 1.

$\Delta S^{\ominus}$ for reaction 1 = ... J K^{-1} mol^{-1} [2]

(c) Group 2 carbonates decompose on heating. The decomposition for one of the Group 2 carbonates, MCO$_3$, is shown in reaction 2.

reaction 2 $MCO_3(s) \rightarrow MO(s) + CO_2(g)$

(i) Predict the sign of the entropy change, ΔS, for reaction 2.

Explain your answer.

.....

..... [1]

(ii) The Gibbs equation is shown.

$$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$$

Fig. 3.1 shows values of the Gibbs free energy change, $\Delta G^{\ominus}$, in kJ mol^{-1}, at different temperatures, T, in K, for reaction 2.

Assume $\Delta H^{\ominus}$ and $\Delta S^{\ominus}$ values for this reaction remain constant over this temperature range.

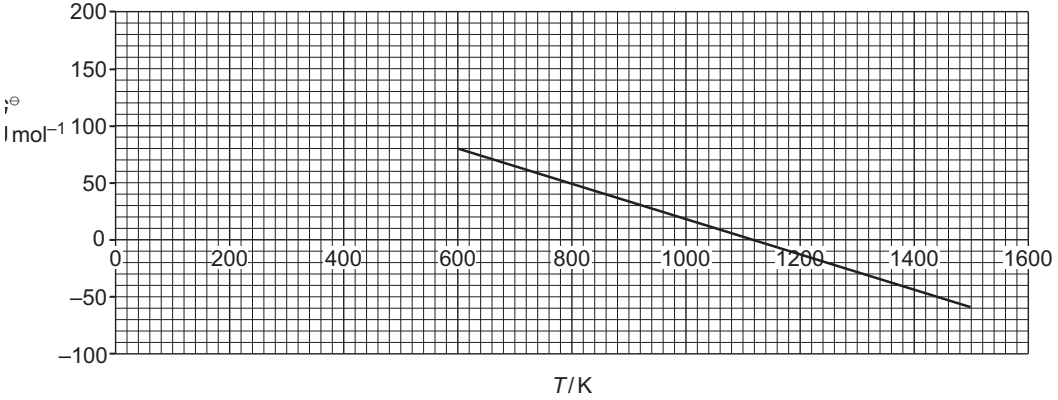


Fig. 3.1

Use the gradient and intercept on the y-axis in Fig. 3.1 and the Gibbs equation to determine:

- $\Delta S^{\ominus}$, in J K^{-1} mol^{-1}, for reaction 2
- the minimum temperature, T, in K, at which the reaction is feasible
- $\Delta H^{\ominus}$, in kJ mol^{-1}, for reaction 2.

$\Delta S^{\ominus}$ for reaction 2 = ... J K^{-1} mol^{-1}

minimum temperature, T = ... K

$\Delta H^{\ominus}$ for r ɔn 2 = ... kJ mol^{-1}
[4]

[Total: 9]