

Exercise walk-through

23.1

Lattice energy and Born-Haber cycles

A-Level Chemistry

Entropy ΔS

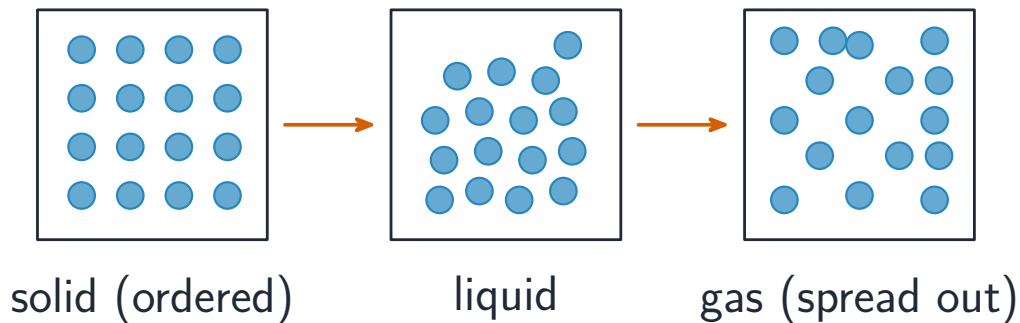
Entropy measures the number of arrangements (“disorder”).

■ solid $\rightarrow$ liquid $\rightarrow$ gas, dissolving: $\Delta S > 0$

■ more gas molecules: $\Delta S > 0$

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

entropy increases: $\Delta S > 0$



Calculating entropy change, and feasibility

from standard entropies

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

the units

$\text{J K}^{-1}\text{mol}^{-1}$ – joules, while ΔH is in kilojoules

the sign

predict it first from the change in moles of gas

Gibbs free energy

$$\Delta G = \Delta H - T\Delta S, \text{ with } T \text{ in kelvin}$$

feasible

when $\Delta G \leq 0$

the feasibility temperature

set $\Delta G = 0$: $T = \frac{\Delta H}{\Delta S}$ – the temperature at which it becomes possible

Convert ΔS from joules to kilojoules before substituting. That factor of 1000 is the single commonest error in this whole topic.

Gibbs free energy ΔG

	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 $\Delta G \leq 0$**

**Gibbs free
energy** :

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

A reaction is **feasible** when $\Delta G \leq 0$. Whether it is can depend on temperature.

ΔG : worked example

$$\Delta H = +120 \text{ kJ mol}^{-1}, \Delta S = +0.200 \text{ kJ K}^{-1} \text{ mol}^{-1}$$

$$\text{at } 298 \text{ K: } \Delta G = 120 - 298 \times 0.200 = +60 \text{ kJ mol}^{-1} \text{ (not yet feasible)}$$

$$\text{Set } \Delta G = 0: T = \frac{\Delta H}{\Delta S} = \frac{120}{0.200} = 600 \text{ K} - \text{feasible above } 600 \text{ K.}$$

Enthalpies of solution and hydration, continued



An instant cold pack feels cold because its salt dissolves endothermically (a positive ΔH_{sol}), driven by the rise in entropy

Exercises 2.1·2.2

2.1 Define **lattice energy**.

[2]

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

[1]

Exercises 2.1·2.2

2.1 Define **lattice energy**.

[2]

2.1

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

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

[1]

2.2

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

Exercises 3.1-3.2

3.1 The enthalpy change of atomisation is always:

[1]

☐

A negative

B positive

C zero

D equal to the lattice energy

☐

A more exothermic

B less exothermic

C zero

D positive

3.2 Down Group 17, the first electron affinity becomes:

[1]

Exercises 3.1-3.2

3.1 The enthalpy change of atomisation is always:

[1]

☐

A negative

B positive

C zero

D equal to the lattice energy

3.1 B

- Atomisation always requires energy (bonds break), so it is positive

3.2 Down Group 17, the first electron affinity becomes:

[1]

☐

A more exothermic

B less exothermic

C zero

D positive

3.2 B

- EA becomes less exothermic down the group

Exercise 3.3 (a)

3.3 Use these data (kJ mol^{-1}) for NaCl: $\Delta H_f = -411$; $\Delta H_{\text{at}}(\text{Na}) = +107$; $\text{IE}_1(\text{Na}) = +496$; $\Delta H_{\text{at}}(\text{Cl}) = +122$; $\text{EA}_1(\text{Cl}) = -349$.

(a) Calculate the lattice energy of NaCl.

[3]

Exercise 3.3 (a)

3.3 Use these data (kJ mol^{-1}) for NaCl: $\Delta H_f = -411$; $\Delta H_{\text{at}}(\text{Na}) = +107$; $\text{IE}_1(\text{Na}) = +496$; $\Delta H_{\text{at}}(\text{Cl}) = +122$; $\text{EA}_1(\text{Cl}) = -349$.

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

(a)

$$\begin{aligned}\bullet \quad \Delta H_{\text{latt}} &= \Delta H_f - [\Delta H_{\text{at}}(\text{Na}) + \text{IE}_1 + \Delta H_{\text{at}}(\text{Cl}) + \text{EA}_1]; = -411 - [107 + 496 + 122 - 349] \\ &= -411 - 376 = -787 \text{ kJ mol}^{-1}\end{aligned}$$

Exercise 3.3 (b)

3.3 Use these data (kJ mol^{-1}) for NaCl: $\Delta H_f = -411$; $\Delta H_{\text{at}}(\text{Na}) = +107$; $\text{IE}_1(\text{Na}) = +496$; $\Delta H_{\text{at}}(\text{Cl}) = +122$; $\text{EA}_1(\text{Cl}) = -349$.

(b) State why the lattice energy is negative.

[1]

Exercise 3.3 (b)

3.3 Use these data (kJ mol^{-1}) for NaCl: $\Delta H_f = -411$; $\Delta H_{\text{at}}(\text{Na}) = +107$; $\text{IE}_1(\text{Na}) = +496$; $\Delta H_{\text{at}}(\text{Cl}) = +122$; $\text{EA}_1(\text{Cl}) = -349$.

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

(b)

- strong electrostatic attractions form between the oppositely charged ions (energy released)

Exercise 3.4

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

Exercise 3.4

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

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