

Exercise walk-through

23.3

Entropy change, ΔS

A-Level Chemistry

Entropy ΔS

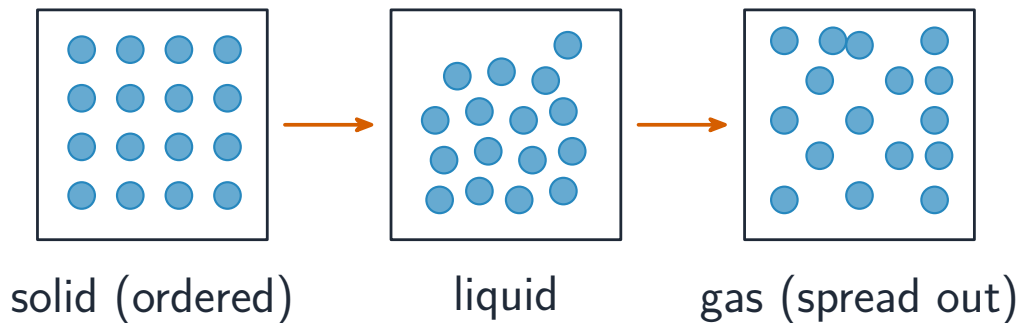
Entropy measures the number of arrangements (“disorder”).

■ 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 2.1·3.1

2.1 Define **entropy**.

[1]

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

[1]

3.1 Which change has the **largest positive** entropy change?

[1]

☐

A freezing of water

B $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

C $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$

D condensing steam

Exercises 2.1·3.1

2.1 Define **entropy**.

[1]

2.1

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

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

[1]

2.2

- negative

3.1 Which change has the **largest positive** entropy change?

[1]

☐

A freezing of water

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D condensing steam

3.1 B

- A solid producing a gas gives the largest increase in entropy

Exercise 3.2

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

[1]

☐

A solid

B liquid

C gas

D crystal

Exercise 3.2

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

[1]

☐

A solid

B liquid

C gas

D crystal

3.2 C

- A gas has the highest entropy

Exercise 3.3 (a)

3.3 For $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, the standard entropies ($\text{J K}^{-1}\text{mol}^{-1}$) are: N_2 192, H_2 131, NH_3 193.

(a) Calculate $\Delta S^\ominus$ for the reaction.

[3]

Exercise 3.3 (a)

3.3 For $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, the standard entropies ($\text{J K}^{-1}\text{mol}^{-1}$) are: N_2 192, H_2 131, NH_3 193.

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

(a)

$$\bullet \Delta S^\ominus = \sum S(\text{products}) - \sum S(\text{reactants}); = 2(193) - [192 + 3(131)] = 386 - 585 = -199 \text{ J K}^{-1}\text{mol}^{-1}$$

Exercise 3.3 (b)

3.3 For $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, the standard entropies ($\text{J K}^{-1}\text{mol}^{-1}$) are: N_2 192, H_2 131, NH_3 193.

(b) Explain the sign of your answer.

[1]

Exercise 3.3 (b)

3.3 For $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, the standard entropies ($\text{J K}^{-1}\text{mol}^{-1}$) are: N_2 192, H_2 131, NH_3 193.

(b) Explain the sign of your answer.

[1]

(b)

- negative – 4 moles of gas become 2 moles, so the disorder decreases

Exercise 3.4

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

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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