

Entropy, Gibbs energy and reaction feasibility

Chemistry grades 11-12 (A Level · IB · Class 11-12) · Worksheet answers

Each answer shows the conversion of units, the substitution and the result with its unit in bold. Where a problem asks you to explain, the key reasoning is given in a sentence so you can compare it with your own wording.

Level: Chemistry grades 11-12 (A Level · IB · Class 11-12) · **Curriculum:** IB R1.4

Problem 1 Which has more entropy?

1. **Liquid water:** particles can move more freely.
2. **CO₂(g):** gases have far more freedom than solids.
3. **NaCl(aq):** separated ions spread through the solution.
4. **Two moles of gas:** more particles give more ways to arrange the energy.

Problem 2 Predicting the sign of ΔS

1. **Positive:** a gas is formed from liquid.
2. **Negative:** 4 mol of gas become 2 mol.
3. **Positive:** a gas is formed from solid.
4. **Negative:** a gas condenses to liquid.

Problem 3 Entropy of alkanes

1. Entropy **increases** as the molecules get larger.
2. $229.6 - 186.3 = \mathbf{+43.3 \text{ J K}^{-1} \text{ mol}^{-1}}$; $270.3 - 229.6 = \mathbf{+40.7 \text{ J K}^{-1} \text{ mol}^{-1}}$.
3. Bigger molecules have more atoms and more ways to vibrate, rotate and bend, so energy can be arranged in more ways.

Problem 4 Gibbs energy of a hand warmer

1. $\Delta S = -120 \div 1000 = \mathbf{-0.120 \text{ kJ K}^{-1} \text{ mol}^{-1}}$.
2. $\Delta G = -85.0 - 298 \times (-0.120) = -85.0 + 35.8 = \mathbf{-49.2 \text{ kJ/mol}}$.
3. **Spontaneous**, because ΔG is negative.

Problem 5 Meaning of the sign of ΔG

1. **Spontaneous.**
2. **Non-spontaneous** (forward direction).
3. **At equilibrium.**
4. The **reverse reaction is spontaneous**, because its ΔG has the opposite sign.

Problem 6 Entropy change for water formation

1. Products: $2 \times 69.9 = 139.8$. Reactants: $2 \times 130.7 + 205.0 = 466.4$.
2. $\Delta S^\circ = 139.8 - 466.4 = -326.6 \text{ J K}^{-1} \text{ mol}^{-1}$.
3. Three moles of gas become a liquid, so the system becomes much more ordered.

Problem 7 Four sign combinations

1. **At all temperatures** (both terms make ΔG negative; ΔG is negative at low T).
2. **Low T only**.
3. **High T only**.
4. **Never**, since ΔH is positive and $-T\Delta S$ is also positive.

Problem 8 Burning methane

1. Products: $213.8 + 2 \times 69.9 = 353.6$. Reactants: $186.3 + 2 \times 205.0 = 596.3$. $\Delta S^\circ = -242.7 \text{ J K}^{-1} \text{ mol}^{-1}$.
2. $\Delta G = -890 - 298 \times (-0.2427) = -890 + 72.3 = -818 \text{ kJ/mol}$.
3. Spontaneous, but it needs a spark because the **activation energy** is high.

Problem 9 Melting point of ice

1. $\Delta G = 0$ at equilibrium between solid and liquid.
2. $T = \Delta H \div \Delta S = 6010 \div 22.0 = 273 \text{ K}$ (0°C).
3. $\Delta G = 6.01 - 290 \times 0.0220 = 6.01 - 6.38 = -0.37 \text{ kJ/mol}$, so **spontaneous**.

Problem 10 Boiling point of ethanol

1. $T = 38\,600 \div 110 = 351 \text{ K}$.
2. $351 - 273 = 78^\circ \text{C}$.
3. Above 351 K the $-T\Delta S$ term outweighs ΔH , so **ΔG becomes negative** and the liquid boils.

Problem 11 Dissolving sodium chloride

1. $\Delta S^\circ = (59.0 + 56.5) - 72.1 = +43.4 \text{ J K}^{-1} \text{ mol}^{-1}$.
2. $\Delta G = 3.9 - 298 \times 0.0434 = 3.9 - 12.9 = -9.0 \text{ kJ/mol}$.
3. The positive entropy change as ions spread through the solution makes $-T\Delta S$ larger than ΔH .

Problem 12 From ΔG to the equilibrium constant

1. $\Delta G^\circ = -8.31 \times 298 \times \ln(4000) = -2476 \times 8.294 = -20\,540 \text{ J} = -20.5 \text{ kJ/mol}$.
2. $\ln K = -15\,000 \div 2476 = -6.057$, so $K = 2.3 \times 10^{-3}$.
3. The **first** reaction (K is greater than 1 and ΔG° is negative).

Problem 13 Reading a ΔG -T graph

1. At $T = 0$, $\Delta G = \Delta H = +60 \text{ kJ/mol}$.
2. Gradient = $-\Delta S$. From $+60$ at 0 K to -60 at 600 K , gradient = $-120 \div 600 = -0.200$, so $\Delta S = +0.200 \text{ kJ K}^{-1} \text{ mol}^{-1}$ ($200 \text{ J K}^{-1} \text{ mol}^{-1}$).
3. $\Delta G = 0$ at **300 K**; the reaction is spontaneous **above 300 K**.

Problem 14 Entropy of the surroundings

1. $\Delta S(\text{surr}) = 1\,367\,000 \div 298 = +4587 \text{ J K}^{-1} \text{ mol}^{-1}$.
2. $\Delta S(\text{total}) = 4587 - 138.4 = +4449 \text{ J K}^{-1} \text{ mol}^{-1}$.
3. $-298 \times 4.449 = -1325.8 \text{ kJ/mol}$; directly, $\Delta G = -1367 - 298 \times (-0.1384) = -1367 + 41.2 = -1326 \text{ kJ/mol}$. They agree.

Problem 15 Decomposition of hydrogen peroxide

1. $\Delta S^\circ = (2 \times 69.9 + 205.0) - 2 \times 109.6 = 344.8 - 219.2 = +125.6 \text{ J K}^{-1} \text{ mol}^{-1}$; $\Delta G = -196 - 298 \times 0.1256 = -233 \text{ kJ/mol}$.
2. ΔG shows only that the reaction is feasible. The **activation energy** is high without a catalyst (such as MnO_2 or an enzyme), so the rate is negligible.

Problem 16 Baking soda decomposition

1. $\Delta S^\circ = (135.0 + 188.8 + 213.8) - 2 \times 101.7 = 537.6 - 203.4 = +334.2 \text{ J K}^{-1} \text{ mol}^{-1}$.
2. $\Delta G = 135.6 - 298 \times 0.3342 = 135.6 - 99.6 = +36.0 \text{ kJ/mol}$: not spontaneous at room temperature.
3. $T = 135\,600 \div 334.2 = 406 \text{ K}$ (133°C).

Problem 17 The contact process

1. $\Delta S^\circ = 2 \times 256.8 - (2 \times 248.2 + 205.0) = 513.6 - 701.4 = -187.8 \text{ J K}^{-1} \text{ mol}^{-1}$.
2. At 298 K : $-198 + 298 \times 0.1878 = -142 \text{ kJ/mol}$. At 1200 K : $-198 + 1200 \times 0.1878 = +27.4 \text{ kJ/mol}$.
3. $T = 198\,000 \div 187.8 = 1054 \text{ K}$.
4. At low T equilibrium favours SO_3 but the rate is too slow; around 700 K with a catalyst is a compromise between yield and rate.

Problem 18 Dinitrogen tetroxide equilibrium

1. $\Delta G = 57.2 - 298 \times 0.1758 = 57.2 - 52.4 = +4.8 \text{ kJ/mol}$.
2. $\ln K = -4812 \div (8.31 \times 298) = -1.943$; $K = 0.143$.
3. $K = 1$ when $\Delta G = 0$: $T = 57\,200 \div 175.8 = 325 \text{ K}$. On heating, the equilibrium shifts to NO_2 and the gas turns **darker brown**.

Problem 19 Three reactions on one graph

1. Line A starts at -120 and falls; line B starts at $+50$ and rises; line C starts at $+90$ and falls steeply, as labelled in the figure.
2. **A** is spontaneous at all temperatures (ΔH negative, ΔS positive); **B** at none (ΔH positive, ΔS negative).
3. $T = 90\,000 \div 200 = \mathbf{450\,K}$; $\Delta G(298\,K) = 90 - 298 \times 0.200 = \mathbf{+30.4\,kJ/mol}$.

Problem 20 Spot the errors

1. ΔH is in kJ but ΔS is in J: the units were not matched, so the $T\Delta S$ term is **1000 times too large**.
2. $\Delta G = -50.0 - 298 \times (-0.100) = -50.0 + 29.8 = \mathbf{-20.2\,kJ/mol}$, spontaneous.
3. $T = \Delta H \div \Delta S = -50\,000 \div -100 = \mathbf{500\,K}$; above 500 K, ΔG is positive.