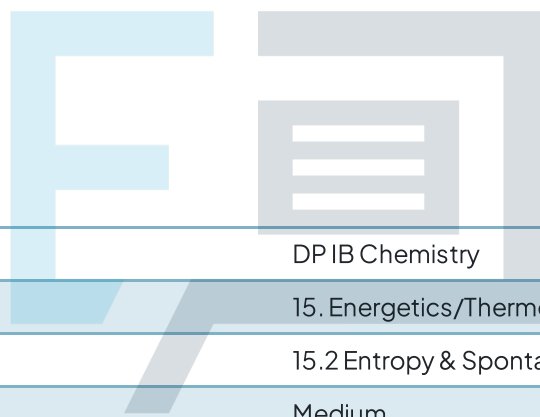




15.2 Entropy & Spontaneity

Mark Schemes



Course	DP IB Chemistry
Section	15. Energetics/Thermochemistry (HL only)
Topic	15.2 Entropy & Spontaneity
Difficulty	Medium

Exam Papers Practice

To be used by all students preparing for DP IB Chemistry HL
Students of other boards may also find this useful



1

The correct answer is **B** because:

- $\Delta G = \Delta H - T\Delta S$
- No matter the value for T (temperature), the value for ΔG will always be greater than zero when ΔH is positive and ΔS is negative, therefore not feasible
- For a reaction to be feasible, the reaction must have a value equal to or lower than zero

If ΔH	and if ΔS	then ΔG is	Feasible?
Is negative	Is positive	Always negative	Always
Is positive	Is negative	Always positive	Never
Is negative	Is negative	Negative at low temperature	Feasible at low temperatures
Is positive	Is positive	Negative at high temperature	Feasible at high temperatures

A is incorrect as	these conditions mean the value for ΔG will always be negative, even if the temperature is low or high. Therefore, the reaction is always feasible.
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C & D are incorrect as	these conditions mean the value for ΔG will always be negative, but the temperature must be as stated, If ΔH is negative and ΔS is negative, the temperature must be low for the reaction to be feasible. If ΔH is positive and ΔS is positive, the temperature must be high for the reaction to be feasible.
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2

The correct answer is **C** because:

- $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$
- So overall:
 - $\Delta G^\ominus = p - 298 \times \frac{q}{1000}$
- **Remember:** The units for ΔH^\ominus and ΔS^\ominus must be the same, so the value for ΔS must be divided by 1000 to convert from $\text{J K}^{-1} \text{mol}^{-1}$ to $\text{kJ K}^{-1} \text{mol}^{-1}$
 - Examiners often comment on units not being converted, this is a common mistake

A is incorrect at	the units for ΔS^\ominus have not been divided by 1000 to convert to $\text{kJ K}^{-1} \text{mol}^{-1}$
B is incorrect as	the incorrect equation has been used to calculate ΔG^\ominus

D is incorrect as	the units for ΔS^\ominus have not been divided by 1000 to convert to $\text{kJ K}^{-1} \text{mol}^{-1}$ and the incorrect equation has been used to calculate ΔG^\ominus
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3

The correct answer is **B** because:

- Statement II is **incorrect**
- You are given the values $\Delta H_r^\ominus = +119 \text{ kJ mol}^{-1}$ and $\Delta S^\ominus = +354.8 \text{ J K}^{-1} \text{ mol}^{-1}$
- ΔH_r^\ominus and ΔS^\ominus are both positive
- The equation to calculate the free energy change is
 - $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$
- At higher temperatures, $-T\Delta S$ will become more and more negative, and will eventually outweigh the effect of ΔH
- The reaction won't be feasible at low temperatures, but if you heat it, there will be a temperature at which it becomes feasible, because ΔG becomes negative
- Such a process is spontaneous at high temperatures and nonspontaneous at low temperatures

If ΔH ...	And if ΔS ...	Then ΔG is...	Feasible...?
Is positive > 0 endothermic	Is positive > 0 More disorder	Negative at high T Positive at low T	Depending on T

- Therefore statement I is correct
 - The reaction will be feasible at high temperatures, but will not be feasible at low temperatures
- Statement III is correct
 - As $\Delta S^\ominus = +354.8 \text{ J K}^{-1} \text{ mol}^{-1}$ is positive, the reaction is becoming more disordered

4

The correct answer is **A** because:

- The equation to calculate Gibbs free energy change is:
 - $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$
 - $\Delta G^\ominus = x - Ty$
- A reaction is not feasible if ΔG^\ominus is greater than 0
- So, in the equation ΔG^\ominus becomes 0
- Therefore the correct rearrangement is:
 - $T = \frac{x}{y}$ becomes $T = \frac{\Delta H^\ominus}{\Delta S^\ominus}$

B, C & D are incorrect as	the Gibbs free energy equation $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$ has not been rearranged correctly in order to calculate T
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5

The correct answer is **D** because:

- In order to determine ΔG^\ominus the required equation is:
 - $\Delta G^\ominus = \sum \Delta G^\ominus_f \text{ products} - \sum \Delta G^\ominus_f \text{ reactants}$
- **Careful:** When calculating ΔG^\ominus you must multiply the values using the correct stoichiometry
 - So, for $6\text{H}_2\text{O}(\text{g})$ you include (6×-228.6) in the equation
- Overall the calculation is:
 - $\Delta G^\ominus = \sum \Delta G^\ominus_f \text{ products} - \sum \Delta G^\ominus_f \text{ reactants}$
 - $\Delta G^\ominus = [(-228.6 \times 6) + (87.6 \times 4)] - (-16.4 \times 4)$
 - This can then continue on to get the final answer - like if this was a question on paper 2:
 - $\Delta G^\ominus = (-1371.6 + 350.4) - (-65.6)$
 - $\Delta G^\ominus = -1021.2 + 65.6$
 - $\Delta G^\ominus = -955.6 \text{ (kJ mol}^{-1}\text{)}$
- Given the ΔG^\ominus_f of oxygen is 0, we do not need to include this in the equation



A is incorrect as	the stoichiometry of the reactants and products has not been used in the calculation
B is incorrect as	$\Delta G^\ominus = \sum \Delta G^\ominus_f \text{ reactants} - \sum \Delta G^\ominus_f \text{ products}$ has been used to calculate ΔG^\ominus
C is incorrect as	the stoichiometry of $\text{H}_2\text{O}(\text{g})$ has not been taken into account



Exam Papers Practice