ENTROPY LIVE CLASS

What's covered in this live class?

Entropy of a system (△S system)
Gibbs Free Energy Calculations
Linking Entropy and Enthalpy

Multiple choice questions
Classic Exam questions
Exam technique

Entropy is a measure of the disorder of a system. It plays a part in determining if a reaction is "feasible" or not. i.e if it will actually occur! It has close links with enthalpy so you really need to have a good handle on that too!
This topic is littered with calculations. They're not difficult but it's easy to make simple errors that can cost you in an exam!

Use the links in the table below to watch the recommended tutorials in the content guide. When you're done, have a go at the questions BEFORE the live class.

AQA	**OCR**	**EDEXCEL**
THERMODYNAMICS (3.1.8) (Entropy & Gibbs Only)	ENTHALPY & ENTROPY (5.2.2)	ENTROPY (13B)

MCQ's

Q1. For which of these changes is there an decrease in entropy?

☐ A Ice melting
☐ B Sodium chloride dissolving
☑ C Condensation of water
☐ D Thermal decomposition of calcium carbonate

Q2. Which is the correct equation for calculating entropy change for a reaction?

☐ A $\Delta S = \sum \Delta S_{reactants} - \sum \Delta S_{products}$
☐ B $\Delta S = \sum \Delta S_{reactants} + \sum \Delta S_{products}$
☑ C $\Delta S = \sum \Delta S_{products} - \sum \Delta S_{reactants}$
☐ D $\Delta S = \sum \Delta S_{products} + \sum \Delta S_{reactants}$

Q3. Which of the following is NOT a factor when calculating ΔG?

☐ A Temperature
☐ B Enthalpy change
☐ C Entropy change
☑ D Temperature change

Q4. Which of the following statements is false?

☐ A ΔG must be zero or below for a reaction to be feasible / spontaneous
☐ B Reactions that have a negative ΔH and a positive ΔS are feasible at any temperature
☐ C Reactions that have a positive ΔH and a negative ΔS are not feasible at any temperature
☑ D Reactions that have a negative ΔH and a negative ΔS are feasible at any temperature

Q5. Which of the following equations is not a true reflection of the relationship between the factors?

☐ A $\Delta G = -T\Delta S + \Delta H$
☐ B $\Delta H = \Delta G + T\Delta S$
☐ C $T = \Delta H / \Delta S$
☑ D $\Delta S = \Delta T / \Delta H$

Classic Exam Questions

Q6. Entropy changes are an important factor in determining the feasibility of reactions.
(a) You are provided with equations for four processes.

A $2SO_{2(g)} + O_2 \rightarrow 2SO_{3(g)}$

B $H_2O_{(l)} \rightarrow H_2O_{(g)}$

C $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$

D $2C_{(s)} + O_{2(g)} \rightarrow 2CO_{(g)}$

For each process, explain why ΔS has the sign shown below.

A: sign of ΔS: negative
reason for sign: fewer gaseous products than reactants.

B: sign of ΔS: positive
reason for sign: Change in state, $(l) \rightarrow (g)$, ∴ ↑ disorder

C: sign of ΔS: negative
reason for sign: Change in state, $(g) \rightarrow (l)$, ∴ ↓ disorder

D: sign of ΔS: positive
reason for sign: Greater moles of (g) in products than reactants.
AND, change in state. No solid products

[4]

(b) Calcium oxide, CaO, is used to make cement. Calcium oxide is manufactured by the thermal decomposition of calcium carbonate.

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)} \quad \Delta H = +178 \text{ kJ.mol}^{-1}$$

Standard entropies of $CaCO_{3(s)}$, $CaO_{(s)}$ and $CO_{2(g)}$ are given in the table below.

substance	$CaCO_3$(s)	CaO(s)	CO_2(g)
***S* / J K^{-1}mol^{-1}**	89	40	214

Using the information in the table, show that the entropy change, ΔS, for the decomposition of calcium carbonate is $0.165 \text{ kJ.K}^{-1}.\text{mol}^{-1}$

Show that calcium carbonate is stable at room temperature (25 °C).

Calculate the minimum temperature needed to decompose calcium carbonate. Show all your working.

$\Delta S = \sum S_{prod} - \sum S_{react}$

$= (40 + 214) - 89 = 165.5 \text{ J.K}^{-}.\text{mol}^{-1}.$ $(\div 1000)$

$\therefore 0.165 \text{ kJ.K}^{-1}.\text{mol}^{-1}.$

$\Delta G = \Delta H - T\Delta S = 178 - (298 \times 0.165) = +128.8$

ΔG is greater than 0, so $CaCO_3$ is stable @ 25°C.

Where ΔG = 0 $\quad T = \frac{\Delta H}{\Delta S} = \frac{+178}{0.165} = 1078.8 \text{ K}$

[7]

Q7. In the Contact Process sulfur dioxide reacts with oxygen to form sulfur trioxide as shown in the equation.

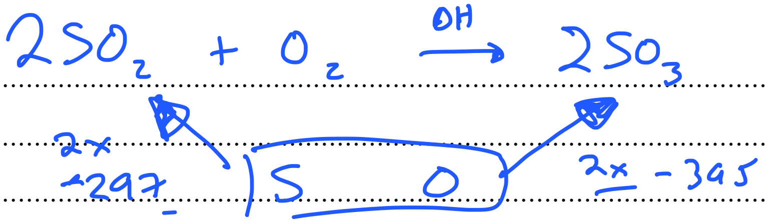
$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

The table shows some thermodynamic data.

	$\Delta H_f^\ominus$ /kJ mol^{-1}	$S^\ominus$/J K^{-1} mol^{-1}
$SO_2(g)$	–297	248
$O_2(g)$	0	205
$SO_3(g)$	–395	256

a) Use data from the table to calculate the standard enthalpy change for this reaction.

[2]



$\Delta H = 2(297) - 2(-395) = \underline{-196\ kJ.mol^{-1}}$

b) Use data from the table to calculate the standard entropy change for this reaction.

[2]

$\Delta S = (2 \times 256) - (2(248) + 205) = -189\ J.K^{-1}.Mol^{-1}$

or $-0.189\ kJ.K^{-1}.mol^{-1}$

c) State what the sign of the entropy change in your answer to part (b) indicates about the product of this reaction relative to the reactants.

[1]

They are more ordered than the reactants.

d) Use your answers to parts (a) and (b) to calculate a value for the free–energy change for this reaction at 50 °C.

(If you were unable to calculate ΔH in part (a) assume a value of −250 kJ.mol⁻¹)

(If you were unable to calculate ΔS in part (b) assume a value of −250 J .K⁻¹.mol⁻¹ These are not the correct values.)

[3]

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

$= -196 - (323 \times -0.189)$

$= -135$ kJ.mol⁻¹ (3sf)

e) Use your answer to part (d) to explain whether the reaction is feasible at 50 °C

[1]

Yes as $\Delta G < 0$.

Q8. Titanium (IV) chloride can be made from titanium (IV) oxide as shown in the equation below:

$$TiO_{2(s)} + 2C_{(s)} + 2Cl_{2(g)} \rightarrow 2CO_{(g)} + TiCl_{4(l)} \qquad \Delta H = -60.0\ kJ.mol^{-1}$$

[3]

Some entropy data are shown the table below:

Substance	TiO_2(s)	C(s)	Cl_2(g)	CO(g)	$TiCl_4$(l)
S° / J K^{-1} mol^{-1}	50.2	5.70	223	198	253

Use the equation and the data in the table to calculate the Gibbs free-energy change for this reaction at 989 °C
Give your answer to the appropriate number of significant figures.
Use your answer to explain whether this reaction is feasible.

[6]

① $\Delta S = \sum S_{prod} - \sum S_{React}$

$= (2\times198 + 253) - (50.2 + 2\times5.7 + 2\times223)$

① $= +141.4\ J.K^{-1}.mol^{-1}$ or ~~+~~ $0.1414\ kJ.K^{-1}.mol^{-1}$ ①

① $\Delta G = \Delta H - T\Delta S$

$= -60.0 - (1262 \times 0.1414) = -238\ kJ.mol^{-1}$ ①

Reaction IS feasible as $\Delta G < 0$ at 989 °C. ①

(3 sf)

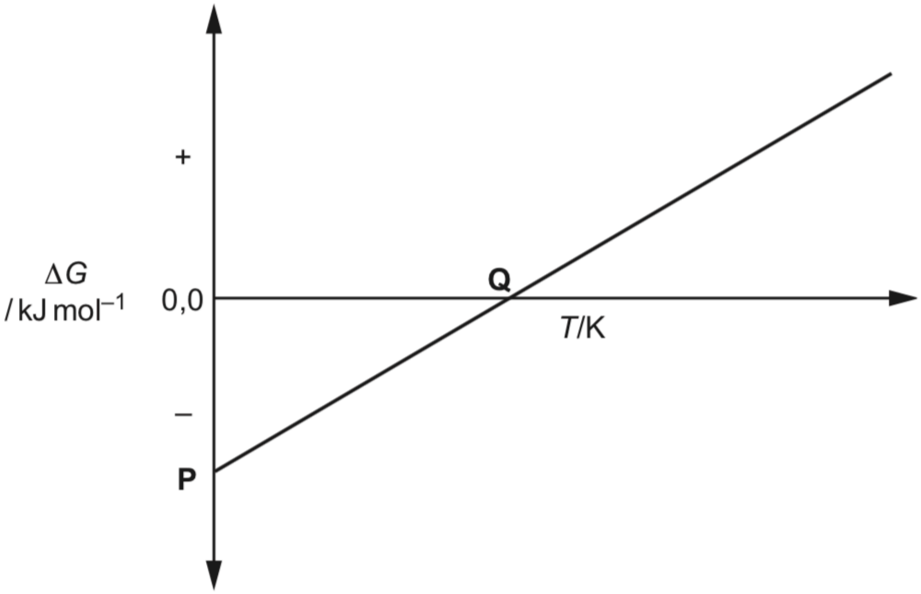
Q9. This question is about free energy changes, ΔG, enthalpy changes, ΔH, and temperature, T.
The Gibbs' equation is shown below.

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

A chemist investigates a reaction to determine how ΔG varies with T.
The results are shown below:

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

$$\equiv \quad y = mx + c$$



What is significant about the gradient of the line and the values P and Q shown in the results?
Explain your reasoning.

[4]

The gradient = $-\Delta S$ ie, it shows the entropy change for the reaction!
As the gradient = $-\Delta S$, As T increases there is a decrease in entropy ie ΔS

$P = \underline{\Delta H}$!

Q = The Temperature, below which the reaction becomes feasible.