InkChemistry

QUESTION WORKBOOK

Entropy and Gibbs free-energy change

Thermodynamics · Unit C08

AQA 3.1.8.2 · OCR A 5.2.2

18 questions · 51 marks

PRINT EDITION

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SECTION A · Short, 1 to 2 marks

A1 EG-Q1 State what is meant by entropy, and give its units. [2]

A2 EG-Q3 Identify which of these reactions has the largest positive entropy change, and give a reason: A, $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$; B, $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$; C, $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$. [2]

A3 EG-Q11 Explain why a graph of ΔG against T for a reaction changes gradient at the boiling point of one of the reactants. [2]

A4 EG-Q14 Suggest why a value of ΔG at 1500 K calculated from standard values of ΔH and ΔS measured at 298 K may be inaccurate. [2]

A5 EG-Q2 State the condition for a reaction to be feasible in terms of ΔG. [1]

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SECTION B · Standard, 3 to 5 marks

B1 EG-Q12

Explain, using $\Delta G = \Delta H - T\Delta S$, why the endothermic decomposition of calcium carbonate is not feasible at room temperature but is feasible at high temperature. [4]

B2 EG-Q16

Evaluate this claim about the Haber process: 'ΔG is positive at the temperature the plant uses, so no ammonia can form there, and the process should be run at room temperature, where ΔG is negative.' [4]

B3 EG-Q18

Calculate the entropy change of the surroundings and the total entropy change at 298 K for $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$, for which $\Delta H = +117$ kJ mol^{-1} and $\Delta S(\text{system}) = +175$ J K^{-1} mol^{-1}. [4]

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B4 EG-Q5 Calculate ΔG at 298 K for $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, for which ΔH = -572 kJ mol^{-1} and ΔS = -327 J K^{-1} mol^{-1}. [3]

B5 EG-Q9 Explain why the entropy of a substance increases far more when it boils than when it melts. [3]

B6 EG-Q10 Explain why diamond does not turn into graphite at room temperature, although ΔG for the change is negative at 298 K. [3]

B7 EG-Q13 Suggest why solid ammonium chloride heated at the bottom of a test tube decomposes into ammonia and hydrogen chloride, $NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)$, but a white solid re-forms on the cool upper part of the tube. [3]

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B8 EG-Q15 Compare the conditions under which a reaction with negative ΔH and negative ΔS is feasible with those for a reaction with positive ΔH and positive ΔS. [3]

B9 EG-Q17 Describe how the entropy of one mole of water changes as it is heated from 0 K to 450 K. [3]

B10 EG-Q19 Explain why an exothermic reaction increases the entropy of its surroundings, and why the increase is larger when the surroundings are cold. [3]

B11 EG-Q20 Calculate the equilibrium constant, K, at 298 K for $2NO_2(g) \rightleftharpoons N_2O_4(g)$, for which $\Delta G = -4.7$ kJ mol^{-1}. Use $\Delta G = -RT \ln K$ and $R = 8.31$ J K^{-1} mol^{-1}. [3]

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B12
EG-Q21

Suggest why a reaction with $\Delta G = +5$ kJ mol^{-1} at 298 K still gives an equilibrium mixture containing a measurable amount of product. $R = 8.31$ J K^{-1} mol^{-1}. [3]

B13
EG-Q23

Show that the thermal decomposition $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ becomes feasible at about 1110 K. $\Delta H = +178$ kJ mol^{-1} and $\Delta S = +161$ J K^{-1} mol^{-1}, both taken as constant. [3]

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