2.3 Energetics and Thermodynamics: Test Pack

AQA 7405 · Calculator allowed · about 55 minutes

Total Marks

Name: _____ Date: ____ / ____ / _____

Answer ALL questions. Show all your working.

1 This question is about standard enthalpy changes and the direction of energy transfer in exothermic and endothermic reactions.

(a) Define what is meant by the standard enthalpy change of formation of a compound. **(2)**

(b) State two standard conditions that must apply when quoting a standard enthalpy change. **(2)**

(c) Hydrogen reacts with oxygen: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, delta H = -572 kJ/mol. State, with a reason, whether this reaction is exothermic or endothermic. **(2)**

(d) Using ideas about bond breaking and bond making, explain why this reaction is exothermic overall. **(2)**

(Total for Question 1 is 8 marks)

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2 This question uses mean bond enthalpies to calculate the enthalpy change for the catalytic cracking of butane.

(a) Butane can be cracked to give ethane and ethene: $C_4H_{10}(g) \rightarrow C_2H_6(g) + C_2H_4(g)$. Explain why, in a mean bond enthalpy calculation for this reaction, the C-H bonds broken and formed do not need to be included. **(2)**

(b) Mean bond enthalpies (kJ/mol): C-C = 347, C=C = 612. Using these values, calculate the enthalpy change for the cracking reaction, showing the bonds broken and bonds formed, and state what the sign of your answer shows about the reaction. **(5)**

(Total for Question 2 is 7 marks)

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3 The table gives lattice enthalpies (magnitude; all values are exothermic) for four sodium halides:

NaF = -918 kJ/mol
NaCl = -780 kJ/mol
NaBr = -742 kJ/mol
NaI = -705 kJ/mol

(a) Describe the trend in lattice enthalpy shown by these values as the halide ion changes from F^- to I^-. **(1)**

(b) Explain this trend in terms of ionic radius and electrostatic attraction. **(3)**

(c) Magnesium oxide, MgO, has a much greater lattice enthalpy (-3791 kJ/mol) than sodium chloride (-780 kJ/mol), even though the ions in MgO are a similar size to those in NaCl. Explain why. **(2)**

(Total for Question 3 is 6 marks)

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4 This question is about the entropy change and thermodynamic feasibility of the thermal decomposition of sodium hydrogencarbonate.

(a) Sodium hydrogencarbonate decomposes on heating: $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$. Standard entropies, S, at 298 K (J/(K mol)): $NaHCO_3(s) = 102$, $Na_2CO_3(s) = 136$, $H_2O(g) = 189$, $CO_2(g) = 214$. Calculate the standard entropy change, delta S, for this reaction. **(3)**

(b) State, with a reason based on the chemical equation, whether you would expect delta S for this reaction to be positive or negative before doing the calculation. **(2)**

(c) The reaction is endothermic, with delta H = +129 kJ/mol. Using your answer to (a), calculate T delta S at 450 K and hence explain, in terms of delta G, why this reaction is thermodynamically feasible at this temperature even though it is endothermic. **(3)**

(Total for Question 4 is 8 marks)

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5 This question is about the thermodynamic feasibility of extracting zinc by reducing zinc oxide with carbon.

(a) State the condition, in terms of delta G, for a reaction to be thermodynamically feasible (spontaneous), and give the equation linking delta G to delta H and delta S. **(2)**

(b) Zinc oxide can be reduced by carbon: $ZnO(s) + C(s) \rightarrow Zn(l) + CO(g)$. Explain, in terms of the signs and relative sizes of delta H and delta S, why this reaction becomes more feasible at higher temperatures. **(2)**

(c) For this reaction, delta H = +237 kJ/mol and delta S = +197 J/(K mol). Calculate the minimum temperature at which the reaction becomes thermodynamically feasible. **(4)**

(d) Suggest one reason why, in industrial practice, this reaction might be carried out at a temperature somewhat higher than the minimum feasibility temperature found in (c). **(2)**

(Total for Question 5 is 10 marks)

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Mark scheme · 2.3 Energetics and Thermodynamics: Test Pack

Question 1

(a) the enthalpy change when 1 mole of a compound is formed from its constituent elements **B1**

(a) with all reactants and products in their standard states, under standard conditions (100 kPa and a stated temperature, usually 298 K) oe **B1**

(a) Answer: The enthalpy change when 1 mole of a compound is formed from its constituent elements, with all substances in their standard states under standard conditions (100 kPa, a stated temperature, usually 298 K).

(b) pressure of 100 kPa **B1**

(b) a stated temperature (usually 298 K) oe; accept solutions at a concentration of 1 mol dm-3 **B1**

(b) Answer: Pressure of 100 kPa; a stated temperature (usually 298 K). Solutions, if present, at a concentration of 1 mol dm-3.

(c) exothermic **B1**

(c) because delta H is negative, so energy is transferred to the surroundings (energy of products is less than energy of reactants) oe **B1**

(c) Answer: Exothermic, because delta H is negative (energy is released to the surroundings).

(d) breaking bonds in the reactants requires an input of energy (endothermic process) oe **M1**

(d) forming new bonds in the products releases energy (exothermic process), and more energy is released forming the new bonds than is absorbed breaking the original bonds oe **M1**

(d) Answer: Breaking bonds in H_2 and O_2 absorbs energy; forming O-H bonds in H_2O releases energy. Because more energy is released forming the new bonds than is absorbed breaking the old ones, the overall reaction is exothermic.

Question 2

(a) butane (C_4H_{10}) contains 10 C-H bonds **B1**

(a) ethane (C_2H_6) and ethene (C_2H_4) together also contain 10 C-H bonds (6 + 4), so exactly the same number of C-H bonds is broken as is reformed, giving no net contribution to delta H, oe **B1**

(a) Answer: The same number of C-H bonds (10) appear on both sides, so they cancel.

(b) bonds broken: 3 mol C-C bonds in butane **M1**

(b) energy to break bonds: 3 x 347 = 1041 kJ/mol **M1**

(b) bonds formed: 1 mol C-C (in ethane) + 1 mol C=C (in ethene); energy released forming bonds: 347 + 612 = 959 kJ/mol **M1**

(b) delta H = bonds broken - bonds formed = 1041 - 959 = +82 kJ/mol, cao **A1**

(b) the positive value shows the reaction is endothermic, consistent with cracking needing heat (and a catalyst) to proceed, oe **B1**

(b) Answer: +82 kJ/mol (endothermic)

Question 3

(a) lattice enthalpy becomes less exothermic (decreases in magnitude) from NaF to NaI oe **B1**

(a) Answer: Lattice enthalpy becomes less exothermic (decreases in magnitude) going from NaF to NaI.

(b) ionic radius of the halide ion increases from F^- to I^- (going down the group) **M1**

(b) as ionic radius increases, the distance between the centres of the oppositely charged ions (Na^+ and the halide ion) increases **M1**

(b) this increased distance weakens the electrostatic force of attraction between the ions, resulting in a lower (less exothermic) lattice enthalpy oe **M1**

(b) Answer: Ionic radius increases from F^- to I^-, increasing the distance between ion centres; this weakens the electrostatic attraction between the ions, giving a smaller (less exothermic) lattice enthalpy.

(c) the ions in MgO carry a greater ionic charge (2+ and 2-) than those in NaCl (1+ and 1-) **M1**

(c) greater ionic charge results in a much stronger electrostatic force of attraction between the ions, increasing the lattice enthalpy oe **M1**

(c) Answer: MgO ions have greater charge (2+ and 2-) than NaCl ions (1+ and 1-); greater charge gives a much stronger electrostatic attraction, and so a much larger (more exothermic) lattice enthalpy.

Question 4

(a) sum of products' entropies: 136 + 189 + 214 = 539 J/(K mol) **M1**

(a) sum of reactants' entropies: 2 x 102 = 204 J/(K mol) **M1**

(a) delta S = 539 - 204 = +335 J/(K mol), cao **A1**

(a) Answer: +335 J/(K mol)

(b) positive **B1**

(b) because gases ($H_2O(g)$ and $CO_2(g)$) are produced from a solid reactant, and gases have much greater entropy (more disordered, more ways of arranging particles and energy) than solids, oe **B1**

(b) Answer: Positive, because gaseous products are formed from a solid reactant.

(c) correct substitution: T delta S = 450 x 0.335 **M1**

(c) T delta S = 150.75 kJ/mol (accept 150-151), which exceeds delta H (129 kJ/mol) **A1**

(c) since delta G = delta H - T delta S, and T delta S is now larger than delta H, delta G is negative at 450 K, so the reaction is thermodynamically feasible at this temperature, oe **B1**

(c) Answer: T delta S = 150.75 kJ/mol > delta H, so delta G is negative and the reaction is feasible at 450 K.

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Question 5

(a) delta G is zero or negative (delta G ≤ 0), oe **B1**

(a) delta G = delta H - T delta S **B1**

(a) Answer: delta G ≤ 0; delta G = delta H - T delta S.

(b) the reaction has a positive delta H (endothermic) and a positive delta S (entropy increases, since a gas, CO, is produced from solids), oe **B1**

(b) as T increases, the (positive) T delta S term becomes larger and eventually outweighs the positive delta H, making delta G negative, oe **B1**

(b) Answer: As T rises, T delta S grows and eventually exceeds delta H, making delta G negative.

(c) set delta G = 0, so delta H = T delta S **M1**

(c) convert units consistently: delta H = 237000 J/mol **M1**

(c) correct substitution: T = 237000/197 **M1**

(c) 1200 K (accept 1200-1210 K), cao **A1**

(c) Answer: Approximately 1200 K

(d) identify a valid practical reason, e.g. to increase the rate of reaction to an economically useful rate (thermodynamic feasibility does not guarantee a fast rate), oe **B1**

(d) elaborate/justify, e.g. operating well above the minimum feasibility temperature also allows for heat losses from the furnace, ensuring the mixture does not fall below the temperature needed to remain feasible, oe **B1**

(d) Answer: A higher temperature ensures a useful reaction rate and allows a margin for heat losses.

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