MEGA LECTURE

IB DP CHEMISTRY

Reactivity 1 · What Drives Chemical Reactions?

TOPICAL WORKSHEET

R1.1 Measuring Enthalpy Changes · R1.2 Energy Cycles in Reactions
R1.3 Energy From Fuels · R1.4 Entropy and Spontaneity

Standard and Higher Level · Syllabus 2025

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Original exam-style questions written for the IB Diploma Programme Chemistry course (first assessment 2025)

Section A — Multiple Choice (15 marks)

Circle the letter of the best answer to each question. Each question is worth 1 mark.

1. The enthalpy change of a reaction, ΔH, is best defined as **[1]**

A the total energy contained within the reactants

B the heat energy exchanged with the surroundings at constant pressure

C the activation energy required to start the reaction

D the entropy change of the surroundings

2. In an exothermic reaction, the sign of ΔH is **[1]**

A positive, and the surroundings warm up

B negative, and the surroundings warm up

C positive, and the surroundings cool down

D negative, and the surroundings cool down

3. In the calorimetry equation q = mcΔT, the symbol c represents **[1]**

A the mass of the solution

B the specific heat capacity of the solution

C the temperature change

D the number of moles of reactant

4. In a calorimetry experiment, the temperature-time graph is extrapolated back to the time of mixing rather than simply reading the maximum temperature. This is done to **[1]**

A make the calculation simpler

B correct for the continuous heat loss (or gain) that occurs during the experiment

C increase the mass of solution used

D account for the specific heat capacity of the calorimeter only

5. Hess's Law states that **[1]**

A the total enthalpy change for a reaction is independent of the route taken, provided initial and final conditions are the same

B enthalpy changes can only be measured directly by calorimetry

C the entropy of the universe always increases

D reactions with a negative ΔH are always spontaneous

6. Hess's Law is valid because enthalpy is **[1]**

A always positive

B a state function, depending only on the initial and final states of a system

C only defined for exothermic reactions

D the same as entropy

7. Using Hess's Law, if step 1 of a two-step reaction pathway has $\Delta H_1 = -110.5$ kJ mol^{-1} and step 2 has $\Delta H_2 = -283.0$ kJ mol^{-1}, the overall enthalpy change for the direct route is **[1]**

A +172.5 kJ mol^{-1}

B +393.5 kJ mol^{-1}

C −393.5 kJ mol^{-1}

D −172.5 kJ mol^{-1}

8. The standard enthalpy change of formation, ΔH_f°, of a compound is defined as the enthalpy change when **[1]**

A one mole of the compound decomposes into its elements

B one mole of the compound is formed from its constituent elements in their standard states

C one mole of the compound is completely combusted in oxygen

D one mole of bonds in the compound is broken

9. Bond enthalpy is best defined as the energy required to **[1]**

A form one mole of a particular covalent bond in the gaseous state

B break one mole of a particular covalent bond in the gaseous state

C break one mole of an ionic bond in the solid state

D vaporize one mole of a liquid

10. In an exothermic reaction, comparing the energy involved in breaking and forming bonds, **[1]**

A more energy is required to break bonds in the reactants than is released forming bonds in the products

B less energy is required to break bonds in the reactants than is released forming bonds in the products

C exactly the same energy is involved in breaking and forming bonds

D no bonds are broken or formed

11. Using bond enthalpies H−H = 436 kJ mol^{-1}, Cl−Cl = 242 kJ mol^{-1}, and H−Cl = 431 kJ mol^{-1}, the enthalpy change for $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ is approximately **[1]**

A +184 kJ mol^{-1}

B −184 kJ mol^{-1}

C +862 kJ mol^{-1}

D −1340 kJ mol^{-1}

12. The enthalpy of combustion of a fuel is defined as the enthalpy change when **[1]**

A one mole of a substance is completely burned in excess oxygen under standard conditions

B one gram of a substance releases heat

C a fuel reacts with a limited supply of oxygen

D one mole of a substance is formed from its elements

13. Entropy, S, is best described as a measure of **[1]**

A the total energy of a system

B the number of ways particles and energy can be arranged (the degree of disorder) in a system

C the heat released by a reaction

D the rate of a chemical reaction

14. For the reaction $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$, the sign of the entropy change of the system, ΔS, is expected to be **[1]**

A positive, since a gas is produced from a solid

B negative, since a gas is produced from a solid

C zero, since mass is conserved

D impossible to determine

15. According to the Gibbs free energy equation, $\Delta G = \Delta H - T\Delta S$, a reaction is thermodynamically spontaneous (feasible) at a given temperature when **[1]**

A $\Delta G < 0$

B $\Delta G > 0$

C $\Delta H < 0$ only, regardless of ΔS

D $\Delta S < 0$ only, regardless of ΔH

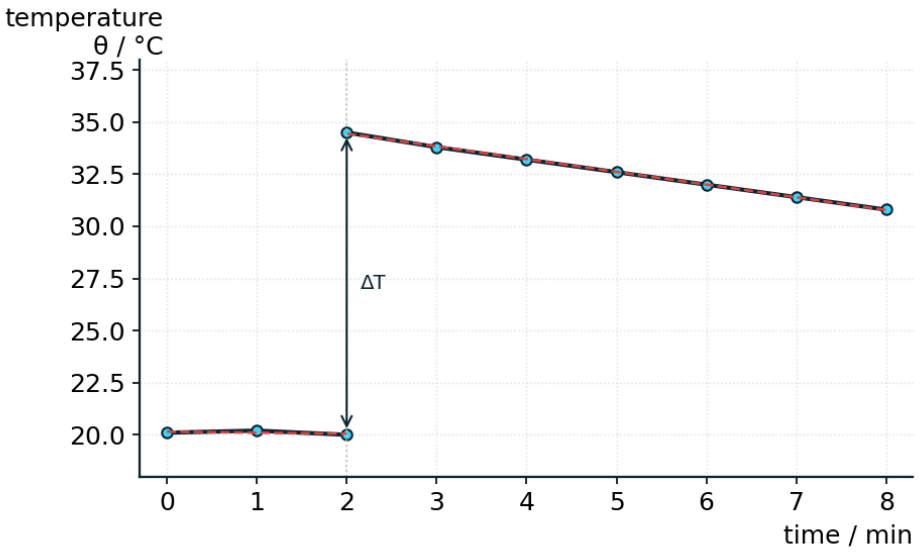
Section B — Structured Questions (66 marks)

Answer all questions in the space provided. Working must be shown for full marks.

R1.1 Measuring Enthalpy Changes

Question 1 [7 marks]

A student adds a solid reactant to 100 cm^3 of dilute acid (density 1.00 g cm^{-3}) in an insulated cup and records the temperature every minute, obtaining the temperature-time graph shown. The reaction released heat and 0.0500 mol of the solid reactant was used (specific heat capacity of the solution, c = 4.18 J g^{-1} K^{-1}).



Temperature-time graph for the calorimetry experiment (given data).

(a) Using the graph, determine the temperature change, ΔT, by extrapolating both lines back to the time of mixing. **[2]**

(b) Calculate the heat energy, q, released by the reaction, using q = mcΔT. **[3]**

(c) Calculate the enthalpy change of the reaction, in kJ mol^{-1}, using the given amount of reactant. **[2]**

Question 2 [6 marks]

Enthalpy changes are usually determined experimentally using calorimetry.

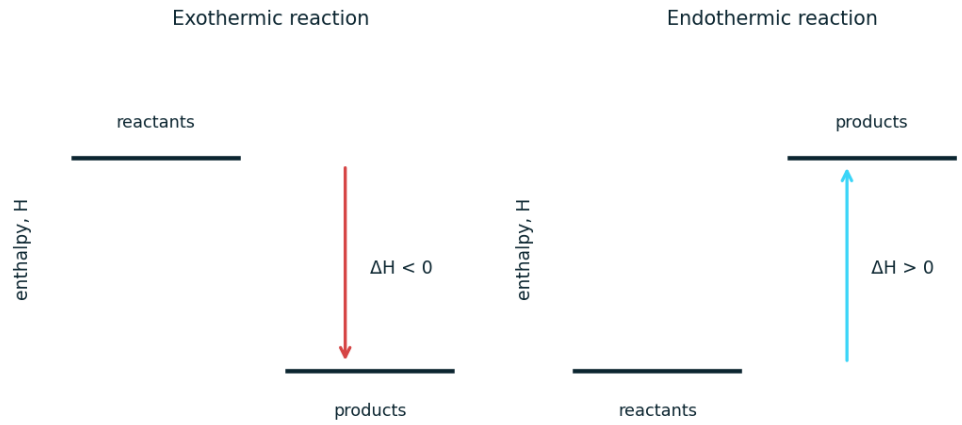
(a) Define the term 'enthalpy change of reaction'. **[2]**

(b) Explain why the temperature-time graph is extrapolated back to the time of mixing, rather than simply using the maximum (or minimum) temperature recorded. **[2]**

(c) Suggest one modification to the experimental procedure that would improve the accuracy of the enthalpy change determined. **[2]**

Question 3 [7 marks]

The diagrams show enthalpy level diagrams for two types of reaction.



Enthalpy level diagrams: exothermic and endothermic reactions (schematic).

(a) Identify which diagram represents an exothermic reaction and which represents an endothermic reaction, and state the sign of ΔH in each case. **[2]**

(b) Explain, in terms of the energy needed to break and form bonds, why the exothermic reaction releases energy overall. **[3]**

(c) Give one example of a chemical or physical process that is endothermic. **[2]**

R1.2 Energy Cycles in Reactions

Question 4 [6 marks]

Hess's Law allows enthalpy changes that are difficult (or impossible) to measure directly to be calculated using an energy cycle.

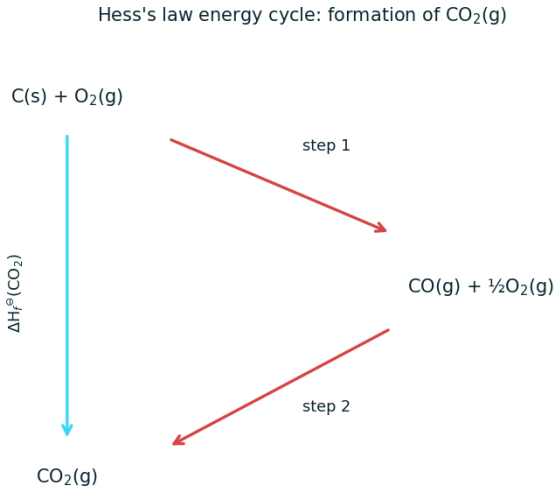
(a) State Hess's Law. **[2]**

(b) Explain why Hess's Law is valid, in terms of the nature of enthalpy. **[2]**

(c) Using standard enthalpies of formation, $\Delta H_f^\circ(CO_2) = -393.5$ kJ mol^{-1} and $\Delta H_f^\circ(H_2O) = -285.8$ kJ mol^{-1} (elements in their standard states have $\Delta H_f^\circ = 0$), calculate ΔH° for the combustion of one mole of carbon: $C(s) + O_2(g) \rightarrow CO_2(g)$. **[2]**

Question 5 [7 marks]

The diagram shows a Hess's law energy cycle for the formation of carbon dioxide, comparing the direct route with a two-step indirect route via carbon monoxide.



Hess's law energy cycle for the formation of $CO_2(g)$ (schematic).

(a) Identify the direct route and the indirect route shown in the diagram. **[2]**

(b) Write an expression, using Hess's Law, relating $\Delta H_f^\circ(CO_2)$ to the enthalpy changes of steps 1 and 2. **[2]**

(c) Given $\Delta H(\text{step 1}) = -110.5\ \text{kJ mol}^{-1}$ and $\Delta H(\text{step 2}) = -283.0\ \text{kJ mol}^{-1}$, calculate $\Delta H_f^\circ(CO_2)$. **[3]**

R1.3 Energy From Fuels

Question 6 [6 marks]

Bond enthalpies can be used to estimate the enthalpy change of a gas-phase reaction. (Bond enthalpies: H−H = 436 kJ mol^{-1}; Cl−Cl = 242 kJ mol^{-1}; H−Cl = 431 kJ mol^{-1}.)

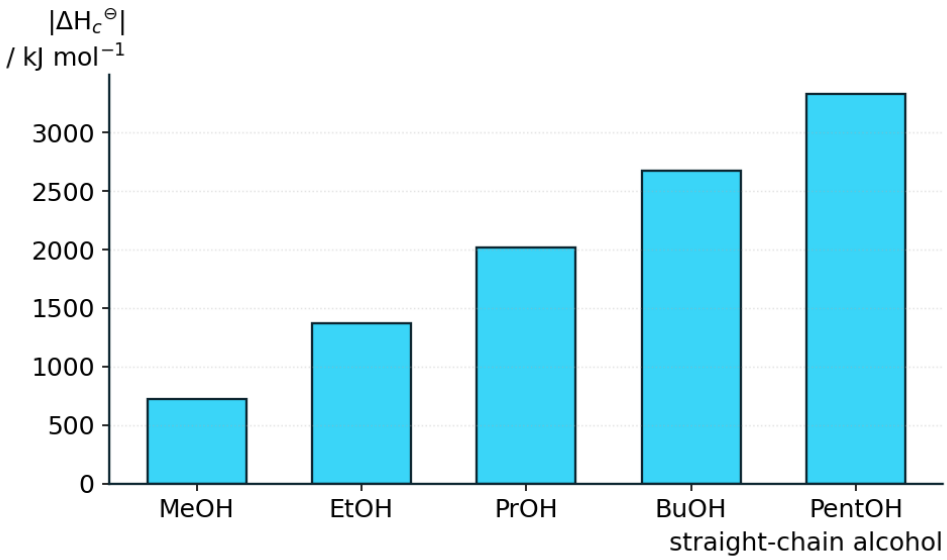
(a) Define the term 'bond enthalpy'. **[2]**

(b) State whether breaking bonds is an exothermic or endothermic process, and whether forming bonds is exothermic or endothermic. **[2]**

(c) Calculate the enthalpy change for the reaction $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$, using the given bond enthalpies. **[2]**

Question 7 [7 marks]

The graph shows the magnitude of the standard enthalpy of combustion for the first five members of the straight-chain, saturated alcohol homologous series.



Enthalpy of combustion of straight-chain alcohols (given data).

(a) Describe the trend shown in the graph as the length of the carbon chain increases. **[2]**

(b) Explain this trend in terms of chemical bonding. **[3]**

(c) The enthalpy of combustion of octane (a major component of petrol/gasoline, M = 114.23 g mol^{-1}) is −5470 kJ mol^{-1}. Using the value for methanol (M = 32.05 g mol^{-1}) from the graph, compare the energy released per gram of each fuel. **[2]**

R1.4 Entropy and Spontaneity

Question 8 [6 marks]

Entropy is a measure of the disorder of a system.

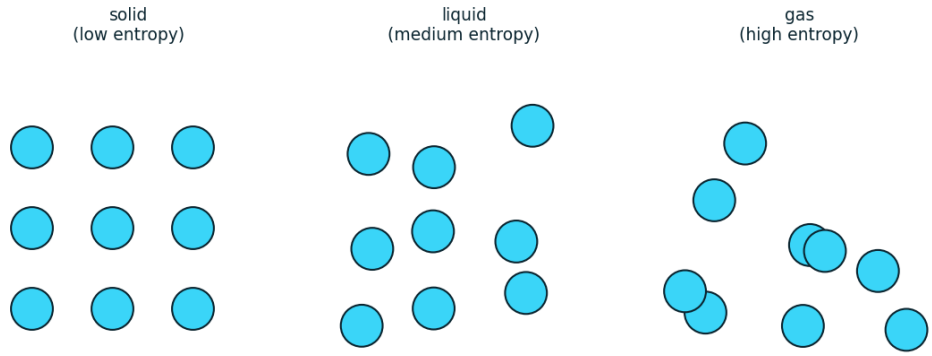
(a) Define the term 'entropy'. **[2]**

(b) Predict the sign of the entropy change, ΔS, for the reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, explaining your reasoning. **[2]**

(c) Explain why the thermal decomposition of a solid to give a gaseous product is associated with a large positive entropy change. **[2]**

Question 9 [7 marks]

The diagrams show a simplified representation of particle arrangement in the solid, liquid, and gas states.



Particle disorder in solid, liquid, and gas states (schematic).

(a) Using the diagrams, rank the three states in order of increasing entropy. **[2]**

(b) Explain, using the diagrams, why the gas state has the highest entropy. **[3]**

(c) Predict the sign of ΔS for the sublimation of solid carbon dioxide (dry ice) directly to CO_2 gas. **[2]**

Question 10 [7 marks]

The spontaneity (thermodynamic feasibility) of a reaction is determined by the sign of the Gibbs free energy change, $\Delta G = \Delta H - T\Delta S$. Consider a reaction with $\Delta H = -92.0$ kJ mol^{-1} and $\Delta S = -198$ J K^{-1} mol^{-1} at $T = 298$ K.

(a) Write the Gibbs free energy equation and state the criterion for a reaction to be thermodynamically spontaneous at a given temperature. **[2]**

(b) Explain why a reaction with a negative ΔH and a negative ΔS can become non-spontaneous at high temperature. **[2]**

(c) Calculate ΔG for this reaction at T = 298 K (convert ΔS to kJ K^{-1} mol^{-1} first), and state whether the reaction is spontaneous at this temperature. **[3]**