13.1 Bond enthalpies and Hess's law

Name: _____ Class: _____ Date: _____ **Total: 16 marks**

KEY WORDS bond enthalpy • Hess's law • energy cycle • enthalpy of formation • Born–Haber cycle – • lattice enthalpy

Objective

Build the skills to answer exam questions on **energy cycles** : bond enthalpies and Hess's law.

You must be able to:

- calculate ΔH from average bond enthalpies and explain the limits of this method
- use Hess's law to find an enthalpy change that cannot be measured directly
- draw and use an energy cycle

1 Key knowledge and worked examples

■ Bond enthalpy and Hess's law

Bond enthalpy: the energy needed to break one mole of a bond in **gaseous** molecules. Values in the data booklet are **averages** over many compounds. Breaking bonds is endothermic; making bonds is exothermic.

$$\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds formed})$$

Hess's law: the enthalpy change of a reaction is the same whichever route is taken, as long as the start and end are the same. Around a cycle, follow the arrows: add a step taken forwards, **subtract** a step taken backwards.

Worked example 1: bond enthalpies. $H_2 + Cl_2 \rightarrow 2HCl$: broken $436+242 = 678$; formed $2 \times 431 = 862$; $\Delta H = 678 - 862 = -184$ **kJ mol**$^{-1}$.

Worked example 2: Hess. If A $\rightarrow$ B is -50 kJ and A $\rightarrow$ C is -80 kJ, then B $\rightarrow$ C is $-80 - (-50) = -30$ **kJ**.

Answer shapes: list bonds with numbers, or draw the cycle and follow the arrows.

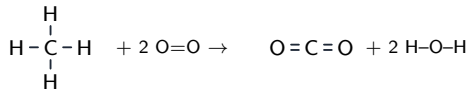
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2 Practice

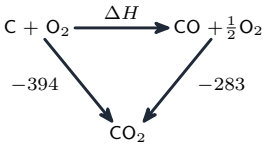
2.1 Explain why ΔH calculated from bond enthalpies is often different from a value measured by calorimetry. [2]

bond	kJ mol^{-1}
C–H	414
O=O	498
O–H	463

2.2 Use the bond enthalpies (kJ mol^{-1}) C–H 414, O=O 498, C=O 804 and O–H 463 to calculate ΔH for $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$. [3]



2.3 Use the cycle to calculate ΔH for $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$. [3]

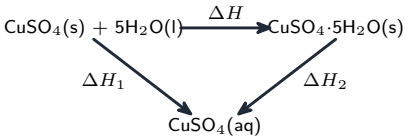


3 Exam-style questions

3.1 The enthalpy change for $CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4{\cdot}5H_2O(s)$ cannot be measured directly. It is found from two enthalpies of solution: anhydrous $CuSO_4$, $\Delta H_1 = -66.5$ kJ mol^{-1}; hydrated $CuSO_4{\cdot}5H_2O$, $\Delta H_2 = +11.7$ kJ mol^{-1}.

(a) State Hess's law. [1]

(b) Calculate the enthalpy change of hydration. [2]



(c) Explain why this enthalpy change cannot be measured directly. [2]

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(d) To find ΔH_1, a student dissolves 0.0200 mol of anhydrous $CuSO_4$ in 50.0 g of water. The temperature rises by 6.2 K. Calculate ΔH_1 from these results. [2]

(e) Suggest why the student's value is less negative than -66.5 kJ mol^{-1}. [1]

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Solutions

Marking: **M1** a correct method, **A1** the correct answer, **B1** an independent mark.

Each answer shows the steps, then the **result**.

2.1

bond enthalpies are averages from many different compounds; they apply to gases, but some substances in the reaction may be liquids

2.2

broken: $4(414) + 2(498) = 2652$; formed: $2(804) + 4(463) = 3460$; $\Delta H = 2652 - 3460 = \mathbf{-808\ kJ\ mol^{-1}}$

2.3

route: $C \rightarrow CO_2$ then backwards $CO_2 \rightarrow CO$; $\Delta H = -394 - (-283) = \mathbf{-111\ kJ\ mol^{-1}}$

3.1

(a) the enthalpy change of a reaction is independent of the route, provided the initial and final states are the same

(b) $\Delta H = \Delta H_1 - \Delta H_2 = -66.5 - 11.7 = \mathbf{-78.2\ kJ\ mol^{-1}}$

(c) it is hard to add exactly the right amount of water; the solid would partly dissolve / the reaction is slow, so the temperature change cannot be measured accurately

(d) $Q = 50.0 \times 4.18 \times 6.2 = 1296$ J; $\Delta H_1 = -\dfrac{1.296}{0.0200} = \mathbf{-64.8\ kJ\ mol^{-1}}$

(e) heat lost to the surroundings

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