



Questions

- 1 Consider the specific heat capacity of the following metals (table 4).

Metal	Specific heat capacity/ $\text{J kg}^{-1} \text{K}^{-1}$
Cu	385
Ag	234
Au	130
Pt	134

▲ Table 4

Which metal will show the greatest temperature increase if 50 J of heat is supplied to a 0.001 kg sample of each metal at the same initial temperature?

- A. Cu C. Au
 B. Ag D. Pt [1]
 IB, May 2007

- 2 When 40 joules of heat are added to a sample of solid H_2O at -16.0°C the temperature increases to -8.0°C . What is the mass of the solid H_2O sample? Specific heat capacity of $\text{H}_2\text{O}(\text{s}) = 2.0 \text{ J g}^{-1} \text{K}^{-1}$

- A. 2.5 g C. 10 g
 B. 5.0 g D. 160 g [1]
 IB, Nov 2007

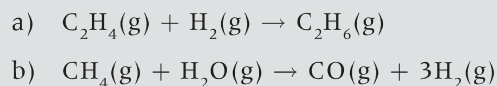
- 3 The temperature of a 2.0 g sample of aluminium increases from 25°C to 30°C . How many joules of heat energy were added? (Specific heat capacity of aluminium = $0.90 \text{ J g}^{-1} \text{K}^{-1}$.)

- A. 0.36 C. 9.0
 B. 2.3 D. 11 [1]
 IB, May 2003

- 4 What is the energy change (in kJ) when the temperature of 20 g of water increases by 10°C ?

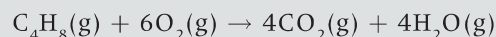
- A. $20 \times 10 \times 4.18$
 B. $20 \times 283 \times 4.18$
 C. $\frac{20 \times 10 \times 4.18}{1000}$
 D. $\frac{20 \times 283 \times 4.18}{1000}$ [1]
 IB, November 2003

- 5 Use data from section 12 of the *Data booklet* to calculate the enthalpy change of reaction for each of these chemical reactions:



- 6 a) Define the term *standard enthalpy change of formation*, ΔH_f^\ominus . [2]

- b) (i) Use the information in table 5 to calculate the enthalpy change for the complete combustion of but-1-ene according to the following equation.



Compound	$\text{C}_4\text{H}_8(\text{g})$	$\text{CO}_2(\text{g})$	$\text{H}_2\text{O}(\text{g})$
$\Delta H_f^\ominus/\text{kJ mol}^{-1}$	+1	-394	-42

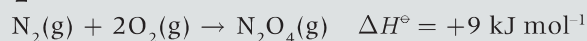
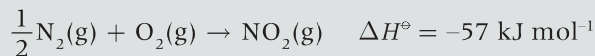
[3]

▲ Table 5

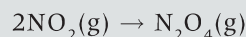
- (ii) Deduce, giving a reason, whether the reactants or the products are more stable. [2]
 (iii) Predict, giving a reason, how the enthalpy change for the complete combustion of but-2-ene would compare with that of but-1-ene based on average bond enthalpies. [1]

IB, May 2007

- 7 The ΔH^\ominus values for the formation of two oxides of nitrogen are given below.



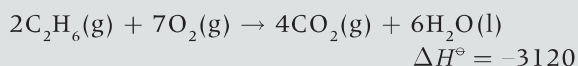
Use these values to calculate ΔH^\ominus for the following reaction (in kJ):



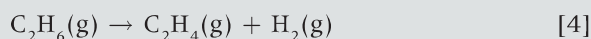
- a) -105 c) +66
 b) -48 d) +123

IB, November 2007

- 8 The standard enthalpy change of three combustion reactions is given below in kJ.



Based on the above information, calculate the standard change in enthalpy, ΔH° , for the following reaction.



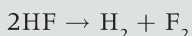
IB, November 2009

- 9 Approximate values of the average bond enthalpies, in kJ mol^{-1} , of three substances are shown in table 6.

H-H	430
F-F	155
H-F	565

▲ Table 6

What is the enthalpy change, in kJ, for this reaction?



- A. +545 B. +20 C. -20 D. -545

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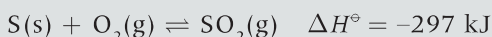
- 10 The reaction between ethene and hydrogen gas is exothermic.

- a) Write an equation for this reaction. [1]
 b) Deduce the relative stabilities and energies of the reactants and products. [2]
 c) Explain, by referring to the bonds in the molecules, why the reaction is exothermic. [2]

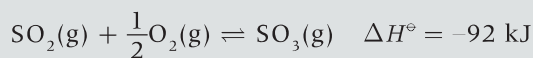
IB, November 2007

- 11 Two reactions occurring in the manufacture of sulfuric acid are shown below:

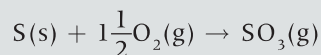
Reaction I



Reaction II

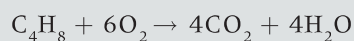


- a) State the name of the term ΔH° . State, with a reason, whether reaction I would be accompanied by a decrease or increase in temperature. [3]
 b) At room temperature sulfur trioxide, SO_3 , is a solid. Deduce, with a reason, whether the ΔH° value would be more negative or less negative if $\text{SO}_3(\text{s})$ instead of $\text{SO}_3(\text{g})$ were formed in reaction II. [2]
 c) Deduce the ΔH° value of this reaction: [1]



IB, November 2005

- 12 But-1-ene gas burns in oxygen to produce carbon dioxide and water vapour according to the following equation.



- a) Use the data in table 7 to calculate the value of ΔH° for the combustion of but-1-ene. [3]

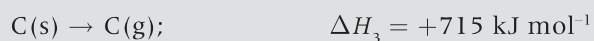
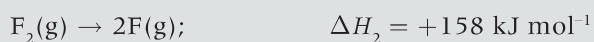
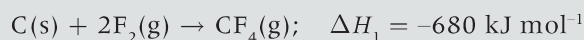
Bond	C-C	C=C	C-H	O=O	C=O	O-H
Average bond enthalpy/ kJ mol^{-1}	348	612	412	496	743	463

▲ Table 7

- b) State and explain whether the reaction above is endothermic or exothermic. [1]

IB, May 2006

- 13 Given the following data:



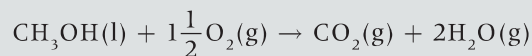
calculate the average bond enthalpy (in kJ mol^{-1}) for the C-F bond.

IB, November 2003

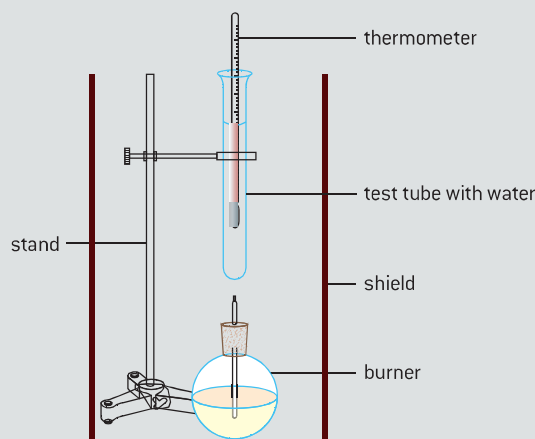


- 14 Methanol is made in large quantities as it is used in the production of polymers and in fuels.

The enthalpy of combustion of methanol can be determined theoretically or experimentally.



- a) Using the information from section 11 of the *Data booklet*, determine the theoretical enthalpy of combustion of methanol. [3]
- b) The enthalpy of combustion of methanol can also be determined experimentally in a school laboratory. A burner containing methanol was weighed and used to heat water in a test tube as illustrated in figure 4.



▲ Figure 4

The data shown in table 8 were collected.

Initial mass of burner and methanol/g	80.557
Final mass of burner and methanol/g	80.034
Mass of water in test tube/g	20.000
Initial temperature of water/°C	21.5
Final temperature of water/°C	26.4

▲ Table 8

- i) Calculate the amount, in mol, of methanol burned. [2]
- ii) Calculate the heat absorbed, in kJ, by the water. [3]
- iii) Determine the enthalpy change, in kJ mol⁻¹, for the combustion of 1 mole of methanol. [2]

- c) The *Data booklet* value for the enthalpy of combustion of methanol is -726 kJ mol⁻¹. Suggest why this value differs from the values calculated in parts a) and b).

- (i) Part a) [1]
- (ii) Part b) [1]

IB, May 2011

- 15 One important property of a rocket fuel mixture is the large volume of gaseous products formed which provide thrust. Hydrazine, N₂H₄, is often used as a rocket fuel. The combustion of hydrazine is represented by the equation below.

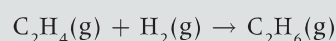


$$\Delta H_c^\ominus = -585 \text{ kJ mol}^{-1}$$

- a) Hydrazine reacts with fluorine to produce nitrogen and hydrogen fluoride, all in the gaseous state. State an equation for the reaction. [2]
- b) Draw the Lewis structures for hydrazine and nitrogen. [2]
- c) Use the average bond enthalpies given in section 11 of the *Data booklet* to determine the enthalpy change for the reaction in part a) above. [3]
- d) Based on your answers to parts a) and c), suggest whether a mixture of hydrazine and fluorine is a better rocket fuel than a mixture of hydrazine and oxygen. [2]

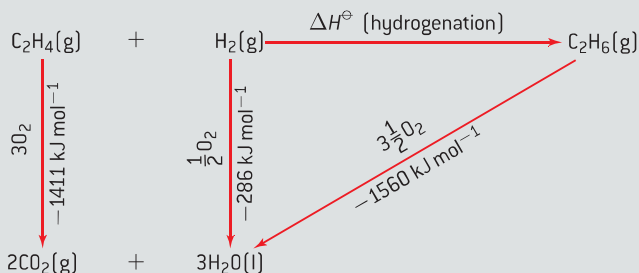
IB, May 2010

- 16 Two students were asked to use information from the *Data booklet* to calculate a value for the enthalpy of hydrogenation of ethene to form ethane.



John used the average bond enthalpies from section 11. Marit used the values of enthalpies of combustion from section 12.

- a) Calculate the value for the enthalpy of hydrogenation of ethene obtained using the average bond enthalpies given in section 11. [2]
- b) Marit arranged the values she found in section 12 into an energy cycle (figure 5).

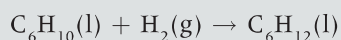


▲ Figure 5

Calculate the value for the enthalpy of hydrogenation of ethene from the energy cycle. [1]

c) Suggest **one** reason why John's answer is slightly less accurate than Marit's answer. [1]

d) John then decided to determine the enthalpy of hydrogenation of cyclohexene to produce cyclohexane.



(i) Use the average bond enthalpies to deduce a value for the enthalpy of hydrogenation of cyclohexene. [1]

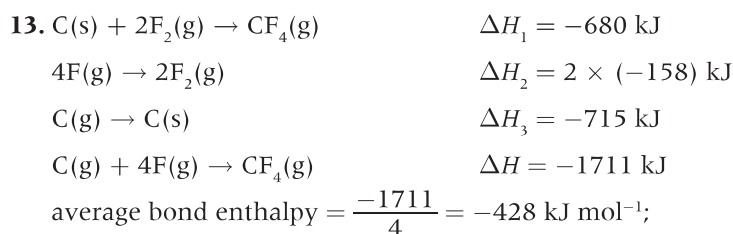
(ii) The percentage difference between these two methods (average bond enthalpies and enthalpies of combustion) is greater for cyclohexene than it was for ethene. John's hypothesis was that it would be the same. Determine why the use of average bond enthalpies is less accurate for the cyclohexene equation shown above, than it was for ethene. Deduce what extra information is needed to provide a more accurate answer. [2]

IB, May 2009



End of topic questions (page 157)

1. C
2. A
3. C
4. C
5. a) $2706 - 2830 = -124$ kJ
 b) $2582 - 2380 = 202$ kJ (NB $\text{C}\equiv\text{O} = 1072$ kJ mol⁻¹)
6. a) the enthalpy change when one mole of compound is formed from its elements in their standard state at standard conditions of 298 K/25°C and 101 325 Pa/1 atm
 b) (i) $\Delta H_p = 4 \times (-242) + 4 \times (-394)$ kJ mol⁻¹
 $\Delta H_R = 1$ kJ mol⁻¹
 $\Delta H^\ominus = (\sum \Delta H_p^\ominus - \sum \Delta H_R^\ominus) = -2545 / -2.55 \times 10^3 / -2550$ kJ mol⁻¹
 (ii) products more stable than reactants; bonds are stronger in products than reactants/
 $H_p < H_R$ /enthalpy/stored energy of products less than reactants;
 (iii) same/equal, because the same bonds are being broken and formed
7. D
8. $\text{C}_2\text{H}_6(\text{g}) + 3\frac{1}{2}\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) \quad \Delta H^\ominus = -1560$ kJ
 $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \quad \Delta H^\ominus = +286$ kJ
 $2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \quad \Delta H^\ominus = +1411$ kJ
 $\text{C}_2\text{H}_6(\text{g}) \rightarrow \text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \quad \Delta H^\ominus = +137$ kJ
9. A
10. a) $\text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$
 b) products more stable than reactants/reactants less stable than products; products lower in energy; reactants higher in energy;
 c) (overall) bonds in reactants weaker and (overall) bonds in product stronger; all bonds in product are σ bonds; weaker π bond broken and a (stronger) σ bond formed; less energy needed to break weaker bonds/more energy produced to make stronger bonds (thus reaction is exothermic);
or
 bond breaking is endothermic (requires energy) and bond making is exothermic (releases energy); stronger bonds in product mean process is exothermic overall;
11. a) standard enthalpy (change) of reaction; temperature increase; reaction is exothermic; sign of ΔH^\ominus is negative;
 b) more negative; heat given out when gas changes to solid; solid has less enthalpy than gas;
 c) -389 kJ
12. a) amount of energy required to break bonds of reactants:
 $8 \times 412 + 2 \times 348 + 612 + 6 \times 496 = 7580$ kJ mol⁻¹
 amount of energy released during bond formation:
 $4 \times 2 \times 743 + 4 \times 2 \times 463 = 9648$ kJ mol⁻¹
 $\Delta H = -2068$ kJ or kJ mol⁻¹
 b) exothermic and ΔH^\ominus is negative; energy is released;



14. a) amount of energy required to break bonds of reactants:

$$3 \times 413 + 358 + 464 + 1.5 \times 498 = 2808 \text{ kJ mol}^{-1}$$

amount of energy released during bond formation of products:

$$4 \times 464 + 2 \times 746 = 3348 \text{ kJ mol}^{-1}$$

$$\Delta H = -540 \text{ kJ mol}^{-1}$$

b) (i) $m(\text{methanol}) = 80.557 - 80.034 = 0.523 \text{ g}$

$$n(\text{methanol}) = \frac{0.523 \text{ g}}{32.05 \text{ g mol}^{-1}} = 0.0163 \text{ mol}$$

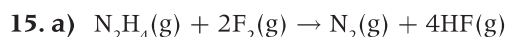
(ii) $\Delta T = 26.4 - 21.5 = 4.9 \text{ K}$;

$$q = mc\Delta T = 20.000 \times 4.18 \times 4.9 \text{ J} = 20.000 \times 4.18 \times 4.9 \times 10^{-3} \text{ kJ} = 0.41 \text{ kJ}$$

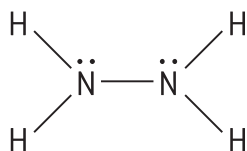
(iii) $\Delta H_c^\ominus = -\frac{0.41 \text{ (kJ)}}{0.0163 \text{ (mol)}} = -25153 \text{ J mol}^{-1} = -25 \text{ kJ mol}^{-1}$

c) (i) bond enthalpies are average values/differ (slightly) from one compound to another (depending on the neighbouring atoms); methanol is liquid, not a gas in the reaction;

(ii) not all heat produced transferred to water; heat lost to surroundings/environment ; incomplete combustion (of methanol); water forms as $H_2O(l)$ instead of $H_2O(g)$;



b) Hydrazine:



Nitrogen:



c) amount of energy required to break bonds of reactants:

$$4 \times 391 + 158 + 2 \times 158 = 2038 \text{ kJ}$$

amount of energy released during bond formation of products:

$$945 + 4 \times 568 = 3217 \text{ kJ}$$

$$\Delta H^\ominus = 2038 - 3217 = -1179 \text{ kJ}$$

d) (N_2H_4/F_2) is a better rocket fuel; 5 vol/mol (g) > 3 vol/mol (g); more moles/greater amount of gas produced; $\Delta H^\ominus (N_2H_4/F_2) > \Delta H^\ominus (N_2H_4/O_2)$ (per mole); (N_2H_4/F_2) reaction more exothermic;

16. a) energy required = $C=C + H-H = 612 + 436$

$$\text{energy released} = C-C + 2(C-H) = 347 + 2 \times 413$$

$$\text{energy required} = C=C + H-H + 4(C-H) = 612 + 436 + 4 \times 413$$

$$\text{energy released} = C-C + 6(C-H) = 347 + 6 \times 413$$

$$\Delta H = (1048 - 1173) = (2700 - 2825) = -125 \text{ kJ mol}^{-1}$$



- b) $\Delta H = -1411 + (-286) - (-1560) = -137 \text{ kJ mol}^{-1}$
- c) the actual values for the specific bonds may be different to the average values; the combustion values referred to the specific compounds;
- d) (i) -125 kJ mol^{-1}
- (ii) average bond enthalpies do not apply to the liquid state; the enthalpy of vaporization/condensation of cyclohexene and cyclohexane.