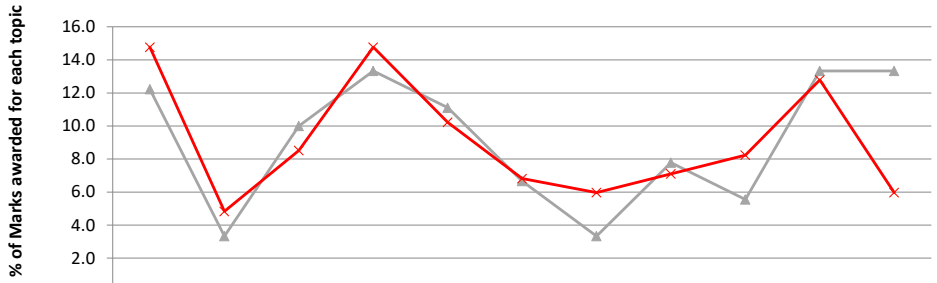
# IB SL 5 EQ Paper 2 Section A & Section B 16w to 99s 233marks

Before 2016 paper 2 was included 4 section B questions, of which you had to chose 2. After 2016 all questions became compulsory on Paper 2

## SL PAPER 1

**Percentage of all marks awarded for each topic from w2012 to w2014, TZ 0, 1 & 2(red crosses) and for P1 just in 2016 s,w and SP (Green)**

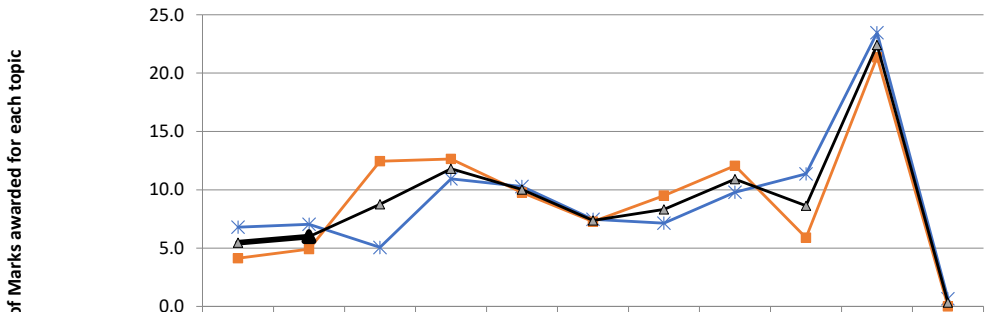



| | Chem 1 | Chem 2 | Chem 3 | Chem 4 | Chem 5 | Chem 6 | Chem 7 | Chem 8 | Chem 9 | Chem 10 | Chem 11 |
|---|---|---|---|---|---|---|---|---|---|---|---|
| P1 2016 | 12.2 | 3.3 | 10.0 | 13.3 | 11.1 | 6.7 | 3.3 | 7.8 | 5.6 | 13.3 | 13.3 |
| P1 '12 to '14 | 14.8 | 4.8 | 8.5 | 14.8 | 10.2 | 6.8 | 6.0 | 7.1 | 8.2 | 12.8 | 6.0 |

**Paper 1 Topic Number**

## STANDARD Level PAPER 2

**Percentage of all marks awarded for each topic from s1999 to w2014 for Paper 2 sections A (compulsory), B (Choose 1 Q out of 3) and A+B. From s2016 onwards, all Paper 2 questions are compulsory**




| | Chem 1 | Chem 2 | Chem 3 | Chem 4 | Chem 5 | Chem 6 | Chem 7 | Chem 8 | Chem 9 | Chem 10 | Chem 11 |
|---|---|---|---|---|---|---|---|---|---|---|---|
| SLP2a&b 08to16 | 6.8 | 7.0 | 5.1 | 10.9 | 10.3 | 7.5 | 7.1 | 9.8 | 11.4 | 23.5 | 0.7 |
| SLP2a&b 99to07 | 4.1 | 4.9 | 12.5 | 12.6 | 9.8 | 7.3 | 9.5 | 12.1 | 5.9 | 21.4 | 0.0 |
| SLP2 All | 5.5 | 6.0 | 8.8 | 11.8 | 10.0 | 7.4 | 8.3 | 10.9 | 8.6 | 22.4 | 0.3 |

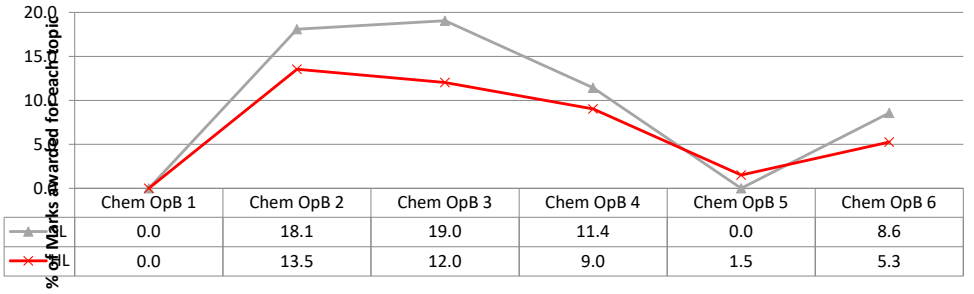
**Paper 2 Topic Number**



## All topics ranked according to their impact on your final grade using exam papers from 1999 to 2016

| | TOPICS and IA | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| | IA | 10 | Option | N.ofSci | 4 | 5 | 8 | 3 | 1 | 9 | 7 | 6 | 2 | 11 |
| Paper 1 | | 12.8 | | | 14.8 | 10.2 | 7.1 | 8.5 | 14.8 | 8.2 | 6.0 | 6.8 | 4.8 | 6.0 |
| Paper 2 16 to 08 | | 23.5 | | | 10.9 | 10.3 | 9.8 | 5.1 | 6.8 | 11.4 | 7.1 | 7.5 | 7.0 | 0.7 |
| Paper 2 07 to 99 | | 21.4 | | | 12.6 | 9.8 | 12.1 | 12.5 | 4.1 | 5.9 | 9.5 | 7.3 | 4.9 | 0.0 |
| Paper 2 ALL | | 22.4 | | | 11.8 | 10.0 | 10.9 | 8.8 | 5.5 | 8.6 | 8.3 | 7.4 | 6.0 | 0.3 |
| Paper 3 | | 0.0 | 57.1 | 42.9 | 0.0 | 0.0 | 0.0 | 0.0 | 0.0 | 0.0 | 0.0 | 0.0 | 0.0 | 0.0 |
| Total % All Marks, Weighted | 20 | 11.5 | 11.4 | 8.6 | 7.7 | 6.1 | 5.8 | 5.2 | 5.1 | 5.1 | 4.5 | 4.3 | 3.4 | 1.3 |
| Rank Order | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 |

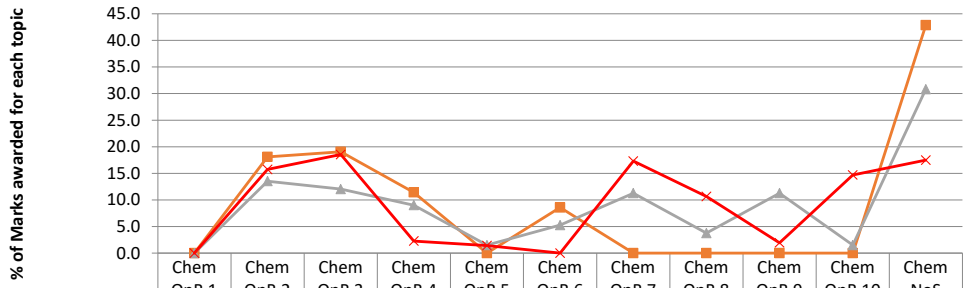
## Human Biochem Mapped to SL Biochemistry for 2016




| | Chem OpB 1 | Chem OpB 2 | Chem OpB 3 | Chem OpB 4 | Chem OpB 5 | Chem OpB 6 |
|---|---|---|---|---|---|---|
| SL | 0.0 | 18.1 | 19.0 | 11.4 | 0.0 | 8.6 |
| HL | 0.0 | 13.5 | 12.0 | 9.0 | 1.5 | 5.3 |

**Biochemistry (2016) Topic Number - X = no longer examined**

## IB SL & HL Human Biochem Option Mapped to Biochemistry Option for 2016

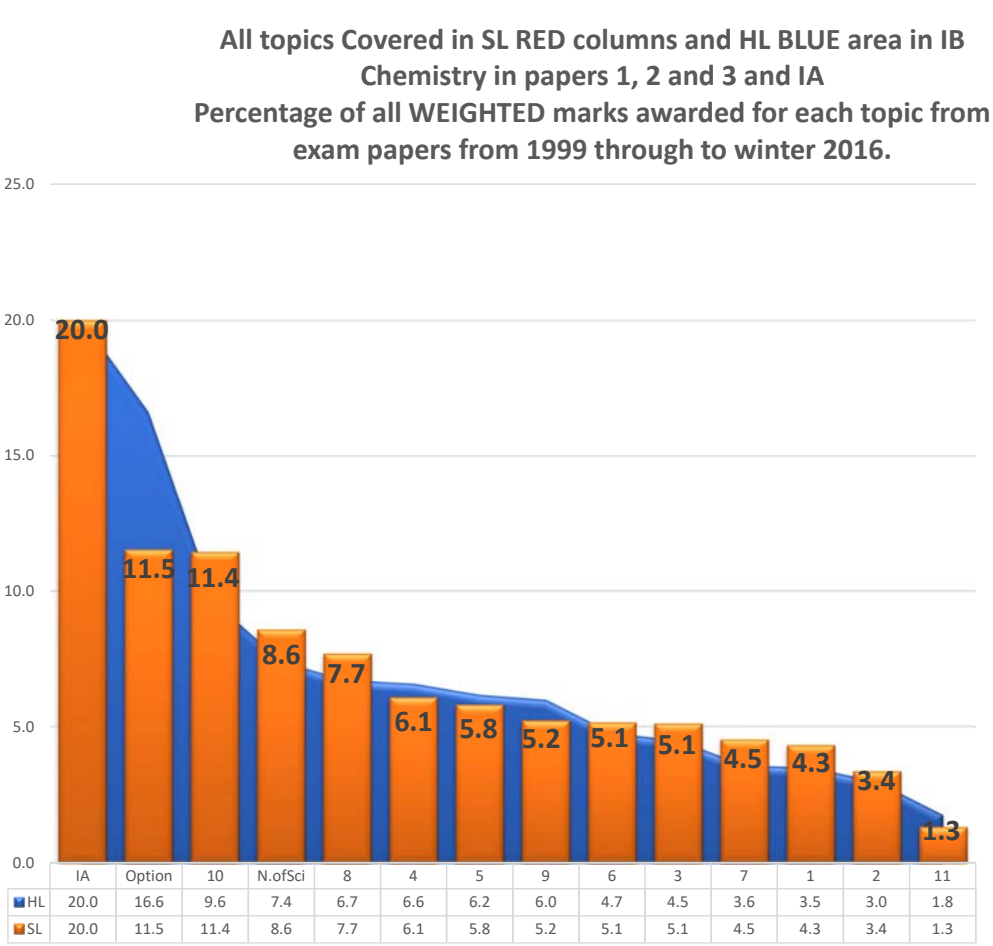



| | Chem OpB 1 | Chem OpB 2 | Chem OpB 3 | Chem OpB 4 | Chem OpB 5 | Chem OpB 6 | Chem OpB 7 | Chem OpB 8 | Chem OpB 9 | Chem OpB 10 | Chem NoS |
|---|---|---|---|---|---|---|---|---|---|---|---|
| SL | 0.0 | 18.1 | 19.0 | 11.4 | 0.0 | 8.6 | 0.0 | 0.0 | 0.0 | 0.0 | 42.9 |
| HL | 0.0 | 13.5 | 12.0 | 9.0 | 1.5 | 5.3 | 11.3 | 3.8 | 11.3 | 1.5 | 30.8 |
| HL & SL Bioch | 0.0 | 15.7 | 18.5 | 2.3 | 1.4 | 0.0 | 17.3 | 10.7 | 1.9 | 14.7 | 17.5 |

**Biochemistry (2016) Topic Number - X = no longer examined**

Essentially, Nature of Science (NoS) is almost half of the paper, sections after 6, so 7, 8 9 & 10 are only found in HL.
This is older data and will be updated in late 2019 (I stopped teaching IB in 2016, teaching A levels instead and started again in the second half of 2019).

## Standard and Higher Level components compared

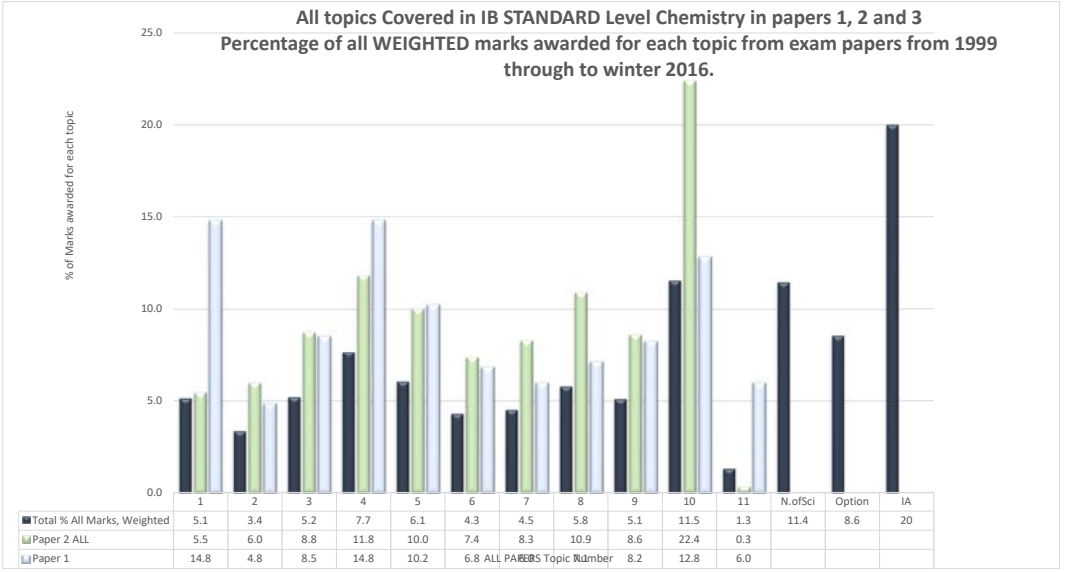



| | IA | Option | 10 | N.ofSci | 8 | 4 | 5 | 9 | 6 | 3 | 7 | 1 | 2 | 11 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| HL | 20.0 | 16.6 | 9.6 | 7.4 | 6.7 | 6.6 | 6.2 | 6.0 | 4.7 | 4.5 | 3.6 | 3.5 | 3.0 | 1.8 |
| SL | 20.0 | 11.5 | 11.4 | 8.6 | 7.7 | 6.1 | 5.8 | 5.2 | 5.1 | 5.1 | 4.5 | 4.3 | 3.4 | 1.3 |

Essentially, IA has the exact same weight, the Option in HL is almost 50% more important than in SL but Topic 10 is more important in SL than HL. All other topics contribute almost equally to a SL and HL grade.




Name




| | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | N.ofSci | Option | IA |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| Total % All Marks, Weighted | 5.1 | 3.4 | 5.2 | 7.7 | 6.1 | 4.3 | 4.5 | 5.8 | 5.1 | 11.5 | 1.3 | 11.4 | 8.6 | 20 |
| Paper 2 ALL | 5.5 | 6.0 | 8.8 | 11.8 | 10.0 | 7.4 | 8.3 | 10.9 | 8.6 | 22.4 | 0.3 | | | |
| Paper 1 | 14.8 | 4.8 | 8.5 | 14.8 | 10.2 | 6.8 | 6.0 | 11.1 | 8.2 | 12.8 | 6.0 | | | |

**The dark blue bars are where your final IB grade will be from:**

1. Your IA is the single most important part of your IB SL, more important than even the Option. Imagine how much time in class, at home and in revision you have or will give to topics 9, 10 and 11. **Your IA, on average, will be worth more to your final grade than all those combined.**
2. The Option is the most important topic for your IB grade compared to the everything else
3. Topic 10, Organic Chemistry, is by far the most important topic for papers 1 and 2.

Name

Q# 1/ IB Chem/2016/w/TZ0/Paper 2 Section A/Standard Level/Q1

**1.** Ethane-1,2-diol, $HOCH_2CH_2OH$, has a wide variety of uses including the removal of ice from aircraft and heat transfer in a solar cell.

(a) Ethane-1,2-diol can be formed according to the following reaction.

$$2CO(g) + 3H_2(g) \rightleftharpoons HOCH_2CH_2OH(g)$$

(iii) Calculate the enthalpy change, $\Delta H^{\ominus}$, in kJ, for this reaction using section 11 of the data booklet. The bond enthalpy of the carbon–oxygen bond in CO (g) is 1077 kJ mol$^{-1}$. [3]

(iv) The enthalpy change, $\Delta H^{\ominus}$, for the following similar reaction is –233.8 kJ.

$$2CO(g) + 3H_2(g) \rightleftharpoons HOCH_2CH_2OH(l)$$

Deduce why this value differs from your answer to (a)(iii). [1]

Q# 2/ IB Chem/2016/SP/TZ0/Paper 2 Section A/Standard Level/Q3

**3.** Many automobile manufacturers are developing vehicles that use hydrogen as a fuel.

(a) Suggest why such vehicles are considered to cause less harm to the environment than those with internal combustion engines. *[1]*



(b) Hydrogen can be produced from the reaction of coke with steam:

$$C(s) + 2H_2O(g) \rightarrow 2H_2(g) + CO_2(g)$$

Using information from section 12 of the data booklet, calculate the change in enthalpy, $\Delta H$, in kJ mol$^{-1}$, for this reaction. *[2]*

Q# 3/ IB Chem/2016/s/TZ0/Paper 2 Section A/Standard Level/Q2

**2.** Impurities cause phosphine to ignite spontaneously in air to form an oxide of phosphorus and water.

(a) (i) 200.0 g of air was heated by the energy from the complete combustion of 1.00 mol phosphine. Calculate the temperature rise using section 1 of the data booklet and the data below. [1]

Standard enthalpy of combustion of phosphine, $\Delta H_c^{\ominus} = -750$ kJ mol$^{-1}$

Specific heat capacity of air = 1.00 J g$^{-1}$ K$^{-1}$ = 1.00 kJ kg$^{-1}$ K$^{-1}$

Q# 4/ IB Chem/2013/s/tz1/Paper 2 Section B/Standard Level/

**7.** To determine the enthalpy change of combustion of methanol, $CH_3OH$, 0.230 g of methanol was combusted in a spirit burner. The heat released increased the temperature of 50.0 cm$^3$ of water from 24.5 °C to 45.8 °C.

(a) (i) Calculate the enthalpy change of combustion of methanol. *[4]*

(ii) Using the theoretical value in Table 12 of the Data Booklet, discuss the experimental results, including **one** improvement that could be made. *[3]*

(b) Methanol can be produced according to the following equation.

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$

Calculate the standard enthalpy change of this reaction using the following data:

I: $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$ $\Delta H^\ominus = -1452\,kJ\,mol^{-1}$

II: $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ $\Delta H^\ominus = -566\,kJ\,mol^{-1}$

III: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H^\ominus = -572\,kJ\,mol^{-1}$ *[3]*

Q# 5/ IB Chem/2012wQ1a

(iv) Using a digital thermometer, the students discovered that the reaction was exothermic. State the sign of the enthalpy change of the reaction, $\Delta H$. *[1]*



Q# 6/ IB Chem/2012/w/TZ0/Paper 2 Section A/Standard Level/Table 10

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

(b) The following equation represents a combustion reaction of propane, $C_3H_8(g)$ when the oxygen supply is limited.

$$C_3H_8(g) + 3\tfrac{1}{2}O_2(g) \rightarrow 3CO(g) + 4H_2O(g)$$

(i) Determine $\Delta H$, the enthalpy change of the reaction, in $kJ\,mol^{-1}$, using average bond enthalpy data from Table 10 of the Data Booklet. The bond enthalpy for the carbon-oxygen bond in carbon monoxide, CO, is 1072 $kJ\,mol^{-1}$. *[3]*

(ii) The CO molecule has dative covalent bonding. Identify a nitrogen-containing positive ion which also has this type of bonding. *[1]*

Q# 7/ IB Chem/2012/s/tz1/Paper 2 Section B/Standard Level/q5Table 10 and 12

(b) Disilane undergoes complete oxidation to form silicon dioxide and water.

$$2Si_2H_6(g)+7O_2(g) \rightarrow 4SiO_2(s)+6H_2O(l)$$

The standard enthalpy change for this reaction, $\Delta H^{\ominus} = -5520$ kJ.

(i) Calculate the standard enthalpy change, in kJ, for the corresponding combustion reaction of 2 moles of ethane, using Table 12 of the Data Booklet. *[1]*

(c) The different properties of the hydrides can be accounted for by the different bond enthalpies of the covalent bonds formed by silicon and carbon.

(i) Define the term *average bond enthalpy*. *[2]*

(ii) Disilane reacts with hydrogen to produce silane, $SiH_4$.

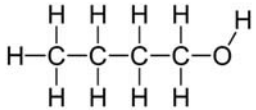
$$Si_2H_6(g)+H_2(g) \rightarrow 2SiH_4(g)$$

Use values from Table 10 of the Data Booklet to calculate the enthalpy change, $\Delta H^{\ominus}$, for this reaction. *[3]*



Q# 8/ IB Chem/2011wQ4

This is the structure of Butan-1-ol:



(b) Deduce the balanced chemical equation for the complete combustion of butan-1-ol. *[1]*

(c) Determine the standard enthalpy change, in kJ mol$^{-1}$, for the complete combustion of butan-1-ol, using the information from Table 10 of the Data Booklet. *[3]*

Q# 9/ IB Chem/2011/w/tz0/Paper 2 Section B/Standard Level/q6

(d) (i) Define the term *endothermic reaction*. *[1]*

(f) Iron, used as the catalyst in the Haber process, has a specific heat capacity of 0.4490 J g$^{-1}$ K$^{-1}$. If 245.0 kJ of heat is supplied to 8.500 kg of iron, initially at a temperature of 15.25 °C, determine its final temperature in K. *[3]*

Q# 10/ IB Chem/2011/w/TZ0/Paper 2 Section A/Standard Level/

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

Q# 11/ IB Chem/2011/s/TZ1/Paper 2 Section A/Standard Level/

**4.** Ethanol is used as a component in fuel for some vehicles. One fuel mixture contains 10% by mass of ethanol in unleaded petrol (gasoline). This mixture is often referred to as Gasohol E10.

(a) Assume that the other 90% by mass of Gasohol E10 is octane. 1.00 kg of this fuel mixture was burned.

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \qquad \Delta H^{\ominus} = -1367 \text{ kJ mol}^{-1}$$

$$C_8H_{18}(l) + 12\tfrac{1}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l) \qquad \Delta H^{\ominus} = -5470 \text{ kJ mol}^{-1}$$

(i) Calculate the mass, in g, of ethanol and octane in 1.00 kg of the fuel mixture. *[1]*

(ii) Calculate the amount, in mol, of ethanol and octane in 1.00 kg of the fuel mixture. *[1]*

(iii) Calculate the total amount of energy, in kJ, released when 1.00 kg of the fuel mixture is completely burned. *[3]*



(b) If the fuel blend was vaporized before combustion, predict whether the amount of energy released would be greater, less or the same. Explain your answer. *[2]*

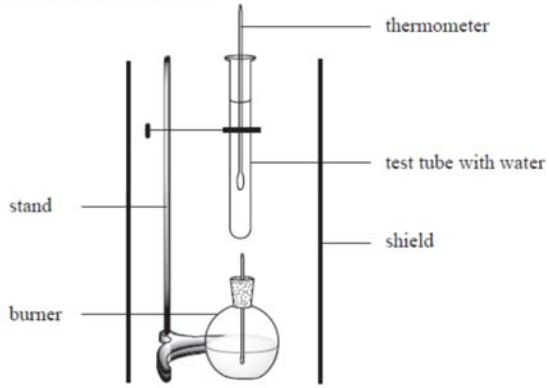
Q# 12/ IB Chem/2011/s/TZ1/Paper 2 Section A/Standard Level/Table 10

**1.** 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.

$$CH_3OH(l) + 1\tfrac{1}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

(a) Using the information from Table 10 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 below.



The following data 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 |

(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]*

Q# 13/ IB Chem/2010/w/tz0/Paper 2 Section B/Standard Level/TABLE 10

**4.** Ethene, $C_2H_4$, and hydrazine, $N_2H_4$, are hydrides of adjacent elements in the periodic table.

(d) Hydrazine is a valuable rocket fuel.

The equation for the reaction between hydrazine and oxygen is given below.

$$N_2H_4(g) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

Use the bond enthalpy values from Table 10 of the Data Booklet to determine the enthalpy change for this reaction. *[3]*

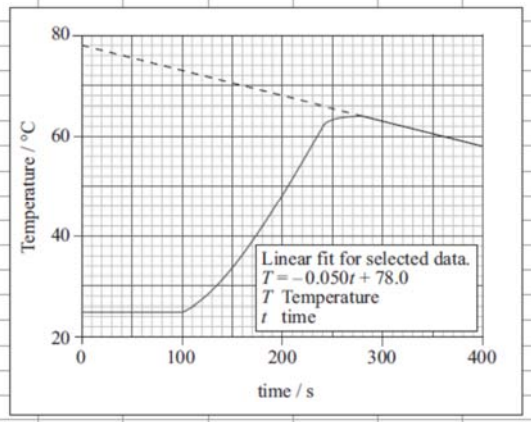
Q# 14/ IB Chem/2010/w/TZ0/Paper 2 Section A/Standard Level/

**1.** The data below are from an experiment to measure the enthalpy change for the reaction of aqueous copper(II) sulfate, $CuSO_4(aq)$ and zinc, $Zn(s)$.

$$Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$$

50.0 $cm^3$ of 1.00 mol $dm^{-3}$ copper(II) sulfate solution was placed in a polystyrene cup and zinc powder was added after 100 seconds. The temperature-time data was taken from a data-logging software program. The table shows the initial 23 readings.

| | A | B | C | D | E | F | G | H |
|---|---|---|---|---|---|---|---|---|
| 1 | **time / s** | **Temperature / °C** | | | | | | |
| 2 | 0.0 | 24.8 | | | | | | |
| 3 | 1.0 | 24.8 | | | | | | |
| 4 | 2.0 | 24.8 | | | | | | |
| 5 | 3.0 | 24.8 | | | | | | |
| 6 | 4.0 | 24.8 | | | | | | |
| 7 | 5.0 | 24.8 | | | | | | |
| 8 | 6.0 | 24.8 | | | | | | |
| 9 | 7.0 | 24.8 | | | | | | |
| 10 | 8.0 | 24.8 | | | | | | |
| 11 | 9.0 | 24.8 | | | | | | |
| 12 | 10.0 | 24.8 | | | | | | |
| 13 | 11.0 | 24.8 | | | | | | |
| 14 | 12.0 | 24.8 | | | | | | |
| 15 | 13.0 | 24.8 | | | | | | |
| 16 | 14.0 | 24.8 | | | | | | |
| 17 | 15.0 | 24.8 | | | | | | |
| 18 | 16.0 | 24.8 | | | | | | |
| 19 | 17.0 | 24.8 | | | | | | |
| 20 | 18.0 | 24.8 | | | | | | |
| 21 | | | | | | | | |
| 22 | | | | | | | | |
| 23 | | | | | | | | |
| 24 | | | | | | | | |




A straight line has been drawn through some of the data points. The equation for this line is given by the data logging software as

$$T = -0.050t + 78.0$$

where $T$ is the Temperature at time $t$.

(a) The heat produced by the reaction can be calculated from the temperature change, $\Delta T$, using the expression below.

Heat change = Volume of $CuSO_4(aq)$ × Specific heat capacity of $H_2O$ × $\Delta T$

Describe **two** assumptions made in using this expression to calculate heat changes. *[2]*

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(b) (i) Use the data presented by the data logging software to deduce the temperature change, $\Delta T$, which would have occurred if the reaction had taken place instantaneously with no heat loss. *[2]*

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(ii) State the assumption made in part (b) (i). *[1]*

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(iii) Calculate the heat, in kJ, produced during the reaction using the expression given in part (a). *[1]*

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(c) The colour of the solution changed from blue to colourless. Deduce the amount, in moles, of zinc which reacted in the polystyrene cup. *[1]*

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. . . . . .

(d) Calculate the enthalpy change, in $kJ\,mol^{-1}$, for this reaction. *[1]*

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(e) An experiment was designed to investigate how the enthalpy change for a displacement reaction relates to the reactivities of the metals involved. The following metals in order of decreasing reactivity were available.

Magnesium *most reactive*
Uranium
Zinc
Iron
Copper
Silver *least reactive*

Excess amounts of each metal were added to 1.00 $mol\,dm^{-3}$ copper(II) sulfate solution. The temperature change was measured and the enthalpy change calculated.

(i) Suggest a possible hypothesis for the relationship between the enthalpy change of the reaction and the reactivity of the metal. *[1]*

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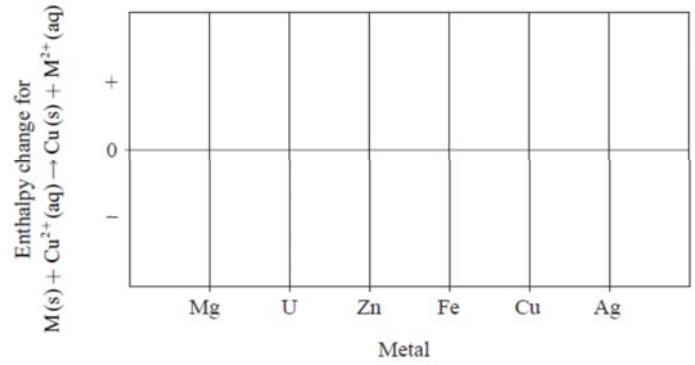
(ii) Sketch a graph on the diagram below to illustrate your hypothesis. *[1]*




Q# 15/ IB Chem/2010/s/tz1/Paper 2 Section B/Standard Level
methanol has the formula $CH_3OH$

**6.** (a) In an experiment to measure the enthalpy change of combustion of ethanol, a student heated a copper calorimeter containing 100 $cm^3$ of water with a spirit lamp and collected the following data.

| | |
|---|---|
| Initial temperature of water: | 20.0 °C |
| Final temperature of water: | 55.0 °C |
| Mass of ethanol burned: | 1.78 g |
| Density of water: | 1.00 g $cm^{-3}$ |

(i) Use the data to calculate the heat evolved when the ethanol was combusted. *[2]*

(ii) Calculate the enthalpy change of combustion per mole of ethanol. *[2]*

(iii) Suggest two reasons why the result is not the same as the value in the Data Booklet. *[2]*

Q# 16/ IB Chem/2009/w/TZ0/Paper 2 Section A/Standard Level/

**3.** The standard enthalpy change of three combustion reactions is given below in kJ.

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \qquad \Delta H^\ominus = -3120$$
$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \qquad \Delta H^\ominus = -572$$
$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \qquad \Delta H^\ominus = -1411$$

Based on the above information, calculate the standard change in enthalpy, $\Delta H^\ominus$, for the following reaction.

$$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g)$$ *[4]*

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Q# 17/ IB Chem/2009/s/TZ1/Paper 2 Section A/Standard Level/Table 10 and 12

**2.** 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.

$$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$$

John used the average bond enthalpies from Table 10. Marit used the values of enthalpies of combustion from Table 12.

(a) Calculate the value for the enthalpy of hydrogenation of ethene obtained using the average bond enthalpies given in Table 10. *[2]*

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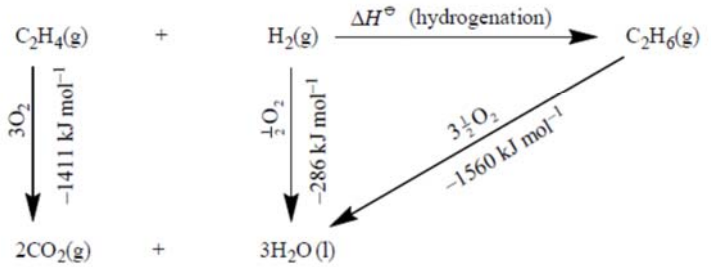
. . . . . .
. . . . . .
. . . . . .

(b) Marit arranged the values she found in Table 12 into an energy cycle.




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

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

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

$$C_6H_{10}(l) + H_2(g) \rightarrow C_6H_{12}(l)$$

(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]*

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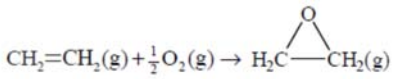
Q# 18/ IB Chem/2008/w/tz0/Paper 2 Section B/Standard Level/Table 10

**4.** (a) (i) Define the term *average bond enthalpy*. *[2]*

(ii) Explain why the H–H bond cannot be used as an example to illustrate average bond enthalpy. *[1]*

(iii) The equation for an important reaction of ethene is given below.



$$CH_2{=}CH_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2C\text{—}CH_2(g)$$

Use information from Table 10 in the Data Booklet to calculate the enthalpy change for this reaction. *[3]*

(iv) Draw a labelled enthalpy level diagram to represent the reaction in part (a) (iii). *[2]*

(b) The standard enthalpy changes for the following reactions can be found in Table 13 of the Data Booklet.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)$$

$$C_8H_{18}(l) + 12\tfrac{1}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l)$$

(i) Use this information to determine the standard enthalpy change for the formation of octane from its elements.

$$8C(s) + 9H_2(g) \rightarrow C_8H_{18}(l)$$ *[4]*

(ii) Predict which of the following reactions has the most negative enthalpy change, and explain your choice.

**I** $H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(g)$

**II** $H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)$ *[2]*



Q# 19/ IB Chem/2008/s/TZ1/Paper 2 Section A/Standard Level/

**3.** (a) The standard enthalpy changes of three combustion reactions at 298 K are given below in kJ mol$^{-1}$.

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta H^\ominus = -394$

$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\quad \Delta H^\ominus = -286$

$CH_3COOH(l) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ $\quad \Delta H^\ominus = -876$

Use the data above to calculate the standard enthalpy change for the following reaction.

$$2C(s) + 2H_2(g) + O_2(g) \rightarrow CH_3COOH(l) \qquad \Delta H^\ominus = ?$$ *[4]*

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Q# 20/ IB Chem/2007/w/tz0/Paper 2 Section B/Standard Level/ not with q5a(i)

**5.** (a) The reaction between ethene and hydrogen gas is exothermic.

$$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$$

(ii) Deduce the relative stabilities and energies of the reactants and products. *[2]*

(iii) Explain, by referring to the bonds in the molecules, why the reaction is exothermic. *[2]*

Q# 21/ IB Chem/2007/s/TZ1/Paper 2 Section A/Standard Level/

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

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(b) Use the information from Table 10 in the Data Booklet to calculate the enthalpy change for the complete combustion of but-1-ene according to the following equation. *[3]*

$$C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$$

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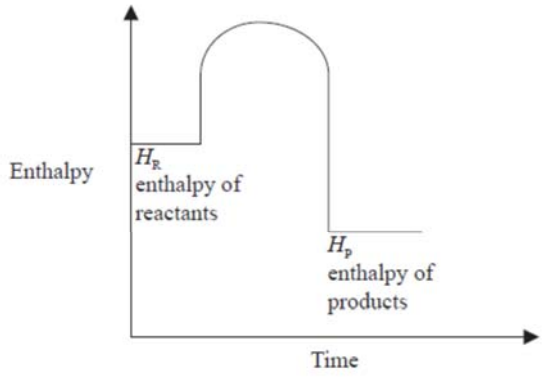
3. (a) Define the term *average bond enthalpy.* [2]

(b) Use the information from Table 10 in the Data Booklet to calculate the enthalpy change for the complete combustion of but-1-ene according to the following equation. [3]

$$C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$$

(c) 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]

(e) The enthalpy level diagram for a certain reaction is shown below.






State and explain the relative stabilities of the reactants and products. [2]

Q# 22/ IB Chem/2007/s/TZ1/Paper 2 Section A/Standard Level/

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

(b) Use the information from Table 10 in the Data Booklet to calculate the enthalpy change for the complete combustion of but-1-ene according to the following equation. [3]

$$C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$$

(c) 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]

Q# 23/ IB Chem/2006/w/TZ0/Paper 2 Section A/Standard Level/TABLE 10

**2.** Methylamine can be manufactured by the following reaction.

$$CH_3OH(g) + NH_3(g) \rightarrow CH_3NH_2(g) + H_2O(g)$$

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

(b) Use information from Table 10 of the Data Booklet to calculate the enthalpy change for this reaction. *[4]*

Q# 24/ IB Chem/2006/s/TZ1/Paper 2 Section A/Standard Level/

**1.** But–1–ene gas, burns in oxygen to produce carbon dioxide and water vapour according to the following equation.

$$C_4H_8 + 6O_2 \rightarrow 4CO_2 + 4H_2O$$

(a) Use the data below to calculate the value of $\Delta H^{\ominus}$ 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 |

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



(c) Calculate the enthalpy change, $\Delta H_4$ for the reaction *[4]*

$$C + 2H_2 + \tfrac{1}{2}O_2 \rightarrow CH_3OH \quad \Delta H_4$$

using Hess's Law and the following information.

$$CH_3OH + 1\tfrac{1}{2}O_2 \rightarrow CO_2 + 2H_2O \qquad \Delta H_1 = -676 \text{ kJ mol}^{-1}$$
$$C + O_2 \rightarrow CO_2 \qquad \Delta H_2 = -394 \text{ kJ mol}^{-1}$$
$$H_2 + \tfrac{1}{2}O_2 \rightarrow H_2O \qquad \Delta H_3 = -242 \text{ kJ mol}^{-1}$$

Q# 25/ IB Chem/2005/w/tz0/Paper 2 Section B/Standard Level/Table 10

**5.** (a) Two reactions occurring in the manufacture of sulfuric acid are shown below:

reaction I $S(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H^{\ominus} = -297$ kJ

reaction II $SO_2(g) + \tfrac{1}{2}O_2(g) \rightleftharpoons SO_3(g)$ $\Delta H^{\ominus} = -92$ kJ

(i) State the name of the term $\Delta H^{\ominus}$. State, with a reason, whether reaction I would be accompanied by a decrease or increase in temperature. *[3]*

(ii) At room temperature sulfur trioxide, $SO_3$, is a solid. Deduce, with a reason, whether the $\Delta H^{\ominus}$ value would be more negative or less negative if $SO_3(s)$ instead of $SO_3(g)$ were formed in reaction II. *[2]*

(iii) Deduce the $\Delta H^{\ominus}$ value of this reaction:

$$S(s) + 1\tfrac{1}{2}O_2(g) \rightarrow SO_3(g) \quad \textit{[1]}$$

(b) (i) Define the term *average bond enthalpy.* *[3]*

(ii) Explain why $Br_2$ is not suitable as an example to illustrate the term *average bond enthalpy.* *[1]*

(iii) Using values from Table 10 of the Data Booklet, calculate the enthalpy change for the following reaction:

$$CH_4(g) + Br_2(g) \rightarrow CH_3Br(g) + HBr(g) \quad \textit{[3]}$$

(iv) Sketch an enthalpy level diagram for the reaction in part (b) (iii). *[2]*

(v) Without carrying out a calculation, suggest, with a reason, how the enthalpy change for the following reaction compares with that of the reaction in part (b) (iii):

$$CH_3Br(g) + Br_2(g) \rightarrow CH_2Br_2(g) + HBr(g)$$ *[2]*

Q# 26/ IB Chem/2005/s/TZ1/Paper 2 Section A/Standard Level/

**1.** In aqueous solution, potassium hydroxide and hydrochloric acid react as follows.

$$KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$$

The data below is from an experiment to determine the enthalpy change of this reaction.

$50.0\ cm^3$ of a $0.500\ mol\ dm^{-3}$ solution of KOH was mixed rapidly in a glass beaker with $50.0\ cm^3$ of a $0.500\ mol\ dm^{-3}$ solution of HCl.
Initial temperature of each solution = 19.6 °C
Final temperature of the mixture = 23.1 °C

(a) State, with a reason, whether the reaction is exothermic or endothermic. *[1]*

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(b) Explain why the solutions were mixed rapidly. *[1]*

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(c) Calculate the enthalpy change of this reaction in $kJ\ mol^{-1}$. Assume that the specific heat capacity of the solution is the same as that of water. *[4]*

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(d) Identify the **major** source of error in the experimental procedure described above. Explain how it could be minimized. *[2]*

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(e) The experiment was repeated but with an HCl concentration of $0.510\ mol\ dm^{-3}$ instead of $0.500\ mol\ dm^{-3}$. State and explain what the temperature change would be. *[2]*

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Q# 27/ IB Chem/2004/w/TZ0/Paper 2 Section A/Standard Level/Table 10

**3.** (a) Define the term *average bond enthalpy*, illustrating your answer with an equation for methane, $CH_4$. *[3]*

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(b) The equation for the reaction between methane and chlorine is

$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$$

Use the values from Table 10 of the Data Booklet to calculate the enthalpy change for this reaction. *[3]*

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(d) Draw an enthalpy level diagram for this reaction. *[2]*



Q# 28/ IB Chem/2004/s/tz1/Paper 2 Section B/Standard Level/table 10

**5.** (a) The standard enthalpy change of formation of $Al_2O_3(s)$ is $-1669\ kJ\ mol^{-1}$ and the standard enthalpy change of formation of $Fe_2O_3(s)$ is $-822\ kJ\ mol^{-1}$.

(i) Use these values to calculate $\Delta H^\ominus$ for the following reaction.

$$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(s) + Al_2O_3(s)$$

State whether the reaction is exothermic or endothermic. *[3]*

(ii) Draw an enthalpy level diagram to represent this reaction. State the conditions under which standard enthalpy changes are measured. *[2]*

(c) Consider the following reaction.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

(i) Use values from Table 10 in the Data Booklet to calculate the enthalpy change, $\Delta H^\ominus$, for this reaction. *[3]*

Q# 29/ IB Chem/2003/w/TZ0/Paper 2 Section A/Standard Level/

**1.** (a) Given the following data:

$$C(s) + 2F_2(g) \rightarrow CF_4(g);\ \Delta H_1 = -680\ kJ\ mol^{-1}$$
$$F_2(g) \rightarrow 2F(g);\ \Delta H_2 = +158\ kJ\ mol^{-1}$$
$$C(s) \rightarrow C(g);\ \Delta H_3 = +715\ kJ\ mol^{-1}$$

calculate the average bond enthalpy (in $kJ\ mol^{-1}$) for the C—F bond. *[4]*

Q# 30/ IB Chem/2001/s/TZ1/Paper 2 Section A/Standard Level/

**1.** (a) Hess's law states that, whether a reaction occurs in one or several steps, the total enthalpy change is the same. Illustrate your understanding of this law by using the data below to calculate the enthalpy change ($\Delta H$) when one mole of solid carbon is converted into carbon monoxide.

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ kJ$$
$$CO(g) + \tfrac{1}{2}O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -283.0\ kJ$$

*[3]*



(b) State what is meant by the term *endothermic reaction*. *[1]*

(c) Enthalpy changes may also be calculated by using bond enthalpies, some values of which ($kJ\ mol^{-1}$) are provided below:

C═C 612; C—H 412; O—H 463; C═O 743; O═O 496.

The balanced equation for the complete combustion of one mole of ethene, $C_2H_4$, in oxygen is shown below:

$$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$$

(i) Use the equation and the bond enthalpy data above to calculate the enthalpy change for the complete combustion of one mole of ethene. *[4]*

(ii) State, with a brief explanation, whether the reaction is endothermic or exothermic. *[1]*

Q# 31/ IB Chem/2000/s/tz1/Paper 2 Section B/Standard Level/

**5.** (a) (i) Explain what is meant by the term *standard enthalpy change of reaction*. *[3]*

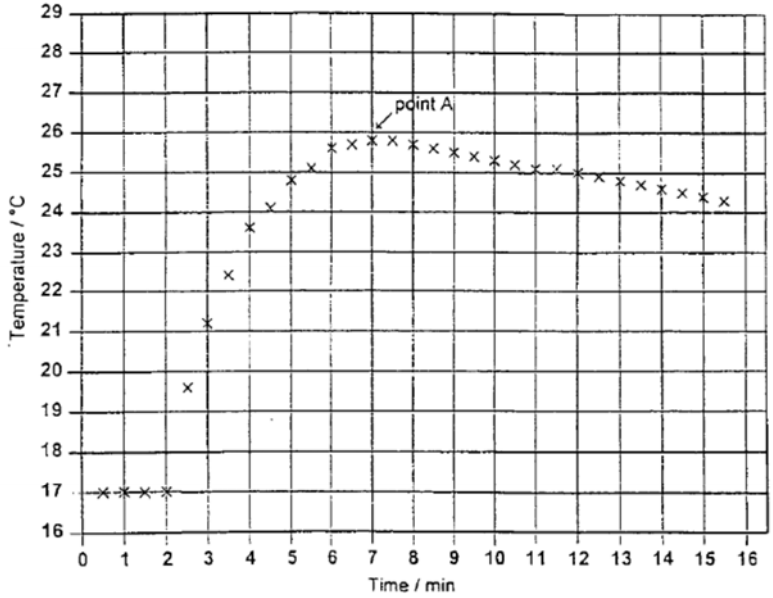
(ii) Describe an experiment to determine the enthalpy change of the reaction between dilute hydrochloric acid and aqueous sodium hydroxide. Show how the value of $\Delta H$ would be calculated from the data obtained. *[9]*

(iii) Draw an enthalpy level diagram for the neutralisation reaction above. Indicate on your diagram the enthalpy change of the reaction and hence compare the relative stabilities of reactants and products. *[4]*

(b) Explain, giving **one** example, the usefulness of Hess's Law in determining $\Delta H$ values. *[4]*

Q# 32/ IB Chem/1999/w/TZ0/Paper 2 Section A/Standard Level/

1. In order to determine the enthalpy change of reaction between zinc and copper(II) sulfate, a student placed 50.0 $cm^3$ of 0.200 $mol\,dm^{-3}$ copper(II) sulfate solution in a polystyrene beaker. The temperature was recorded every 30 seconds. After two minutes 1.20 g of powered zinc was added. The solution was stirred and the temperature recorded every half minute for the next 14 minutes. The results obtained were then plotted to give the following graph:




(a) Write the equation for the reaction taking place. *[1]*

.....

(b) Determine which of the two reagents was present in excess. *[2]*

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.....

(c) The highest temperature is reached at point **A**. Explain what is happening in the system at this point. *[1]*

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(d) By drawing a suitable line on the graph estimate what the rise in temperature would have been if the reaction had taken place instantaneously. *[2]*

.....



(e) Calculate how much heat was evolved during the reaction. Give your answer to **three** significant figures. *[2]*

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(f) What is the enthalpy change of reaction in $kJ\,mol^{-1}$? *[1]*

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(g) The accepted value for the enthalpy change of reaction is $-218\ kJ\,mol^{-1}$. What is the percentage error for the value obtained in this experiment? *[1]*

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(h) Suggest **one** reason why there is disagreement between the experimental value and the accepted value. *[1]*

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Q# 33/ IB Chem/1999/s/TZ1/Paper 2 Section A/Standard Level/

2. (a) Define *average bond enthalpy*. *[2]*

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(b) Using the following average bond enthalpies ($kJ\,mol^{-1}$);

C−H 412, O=O 496, C=O 743, O−H 463

calculate the enthalpy change for the following reaction: *[2]*

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

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# Mark Scheme IB SL 5 EQ Paper 2 Section A & B 16w to 99s 233marks

Q# 1/ IB Chem/2016/w/TZ0/Paper 2 Section A/Standard Level/Q1

| | | | | | |
|---|---|---|---|---|---|
| 1. | a | iii | *Bonds broken*: 2C≡O + 3(H–H) / 2(1077 kJ mol$^{-1}$) + 3(436 kJ mol$^{-1}$) / 3462 «kJ» ✔<br><br>*Bonds formed*: 2(C–O) + 2(O–H) + 4(C–H) + (C–C) / 2(358 kJ mol$^{-1}$) + 2(463 kJ mol$^{-1}$) + 4(414 kJ mol$^{-1}$) + 346 kJ mol$^{-1}$ / 3644 «kJ» ✔<br><br>«Enthalpy change = bonds broken – bonds formed = 3462 kJ – 3644 kJ =» –182 «kJ» ✔ | *Award [3] for correct final answer.*<br>*Award [2 max] for «+»182 «kJ».* | 3 |
| 1. | a | iv | in (a)(iii) gas is formed and in (a)(iv) liquid is formed<br>*OR*<br>products are in different states<br>*OR*<br>conversion of gas to liquid is exothermic<br>*OR*<br>conversion of liquid to gas is endothermic<br>*OR*<br>enthalpy of vapourisation needs to be taken into account ✔ | *Accept product is «now» a liquid.*<br><br>*Accept answers referring to bond enthalpies being means/averages.* | 1 |

Q# 2/ IB Chem/2016/SP/TZ0/Paper 2 Section A/Standard Level/Q3

| Question | | | Answers | Notes | Total |
|---|---|---|---|---|---|
| 3. | a | | only water/$H_2O$ produced ‹so non-polluting› ✔ | | 1 |
| | b | | $\Delta H$ = [(–393.5)] – [(2)(–241.8)] ✔<br>+90.1 kJ ✔ | *Award [2] for correct final answer.* | 2 |

Q# 3/ IB Chem/2016/s/TZ0/Paper 2 Section A/Standard Level/Q2

| Question | | | Answers | Notes | Total |
|---|---|---|---|---|---|
| 2. | a | i | temperature rise «= $\frac{750 \times 1.00}{0.2000 \times 1.00}$» = 3750 «°C/K» ✔ | *Do not accept –3750.* | 1 |

Q# 4/ IB Chem/2013/s/tz1/Paper 2 Section B/Standard Level/

7. (a) (i) $(q = mc\Delta T =)\ 0.0500 \times 4.18 \times 21.3 = 4.45$ (kJ);
*Do not accept m = 0.05023 kg.*

$(n\ methanol =)\ \frac{0.230}{32.05} = 7.18 \times 10^{-3}$ (mol);

$\Delta H = \frac{4.45}{7.18 \times 10^{-3}}$;

$\Delta H = -6.20 \times 10^{2}$ kJ mol$^{-1}$; **[4]**

*Accept integer values of molar mass.*
*Final answer must have negative sign and correct units.*
*Award [4] for correct final answer with correct units.*

(ii) less heat is liberated than theoretically/–726 kJ mol$^{-1}$;
probably due to heat loss/incomplete combustion;
determine heat capacity of calorimeter and take heat absorbed by calorimeter into account / any suitable insulation method / measure temperature with time and extrapolation of graph to compensate heat loss / *OWTTE*; **[3]**
*If the value calculated in (a) (i) is more exothermic than theoretically, allow ECF for M1 and for improvement if consistent.*

(b) $\Delta H^{\ominus} = \frac{1}{2}$II + III – $\frac{1}{2}$I / correct diagram/energy cycle;
–283 – 572 – (–726);
–129 (kJ mol$^{-1}$); **[3]**
*Award [3] for correct final answer.*

Q# 5/ IB Chem/2012wQ1a

(iii) $\left(\frac{8.60}{10.3} \times 100 =\right)$ 83.5%; **[1]**
*Allow answers in the range of 82.5 to 84.5%.*



Q# 6/ IB Chem/2012/w/TZ0/Paper 2 Section A/Standard Level/

2. (a) energy needed to break (1 mol of) a bond in a gaseous molecule/state/phase;
average calculated from a range of similar compounds / *OWTTE*; **[2]**
*Do not accept similar bonds instead of similar compounds.*
*M2 can be scored independently.*

(b) (i) *Bonds breaking*:
$2 \times (C–C) + 8 \times (C–H) + 3.5 \times (O–O)$
$= (2)(347) + (8)(413) + (3.5)(498)$
$= 5741$ (kJ mol$^{-1}$);

*Bonds forming*:
$3 \times (CO) + 8 \times (O–H)$
$= (3)(1072) + (8)(464) = 6928$ (kJ mol$^{-1}$);

*Enthalpy change*:
$(5741 – 6928 =) –1187$ (kJ mol$^{-1}$); **[3]**
*Award [3] for correct final answer.*

(ii) $NH_4^+$/ammonium / $N_2H_5^+$/hydrazinium / $CH_3NH_3^+$/methylammonium/ methanaminium / $H_2NO_3^+$ /nitrooxonium; **[1]**

Q# 7/ IB Chem/2012/s/tz1/Paper 2 Section B/Standard Level/q5

(b) (i) (–1560 × 2 =) –3120 (kJ); **[1]**

(c) (i) energy needed to break (one mol of) a bond in a gaseous molecule;
averaged over similar compounds; **[2]**
*Do not allow averaged over several compounds.*

(ii) *Bonds broken:*
6Si–H, Si–Si, H–H / (+)2570 (kJ);
*Bonds formed:*
8Si–H / (–)2544 (kJ);
+26 (kJ);

**OR**

*Bonds broken:*
Si–Si, H–H / (+)662 (kJ);
*Bonds formed:*
2Si–H / (–)636 (kJ);
+26 (kJ); **[3]**

Q# 8/ IB Chem/2011wQ4

(b) $CH_3(CH_2)_3OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$; **[1]**
*Allow $C_4H_9OH$ or $C_4H_{10}O$ for $CH_3(CH_2)_3OH$.*
*Ignore state symbols.*

(c) *Bonds broken:*
(6)(O=O) + (3)(C–C) + (1)(O–H) + (1)(C–O) + (9)(C–H) /
((6)(498) + (3)(347) + (1)(464) + (1)(358) + (9)(413) =) 8568 (kJ mol$^{-1}$);

*Bonds formed:*
(8)(C=O) + (10)(O–H) / ((8)(746) + (10)(464) =) 10608 (kJ mol$^{-1}$);

$\Delta H$ = (8568 – 10608 =) – 2040 (kJ mol$^{-1}$); **[3]**
*Award [3] for correct final answer.*
*Award [2] for +2040 (kJ mol$^{-1}$).*

Q# 9/ IB Chem/2011/w/tz0/Paper 2 Section B/Standard Level/q6

(d) (i) (system) absorbs/takes in heat from surroundings / *OWTTE*; **[1]**
*Allow standard enthalpy change/ $\Delta H^{\ominus}$ positive.*
*Allow bond breaking more energetic then bond formation / OWTTE.*
*Absorbs/takes in heat alone not sufficient for mark.*

(f) $q = mc\Delta T = 2.450 \times 10^5 = (8.500 \times 10^3)(0.4490)(T_f - 15.25)$;
$T_f = 79.44\,^{\circ}C / \Delta T = 64.19\,(^{\circ}C / K)$;
$T_f = (79.44 + 273) = 352\,(K)$; **[3]**
*Award **[3]** for correct final answer.*
*Accept the use of 273.15 K instead of 273 K giving final value of 352.59 K.*
*For M1 and M2 award **[1 max]** for use of $q = mc\Delta T$ if incorrect units of m and c are used.*

Q# 10/ IB Chem/2011/w/TZ0/Paper 2 Section A/Standard Level/

**4.** (a) energy required to break (1 mol of) a bond in a gaseous molecule/state;
*Accept energy released when (1 mol of) a bond is formed in a gaseous molecule/ state / enthalpy change when (1 mol of) bonds are formed or broken in the gaseous molecule/state.*
average values obtained from a number of similar bonds/compounds / *OWTTE*; **[2]**

Q# 11/ IB Chem/2011/s/TZ1/Paper 2 Section A/Standard Level/

**4.** (a) (i) (10% 1000 g =) 100 g ethanol **and** (90% 1000 g =) 900 g octane; **[1]**

(ii) $n$(ethanol) = 2.17 mol **and** $n$(octane) = 7.88 mol; **[1]**

(iii) $E_{\text{released from ethanol}} = (2.17 \times 1367) = 2966$ (kJ);
$E_{\text{released from octane}} = (7.88 \times 5470) = 43104$ (kJ);
total energy released = $(2966 + 43104) = 4.61 \times 10^4$ (kJ); **[3]**
*Award **[3]** for correct final answer.*
*Accept answers using whole numbers for molar masses and rounding.*

(b) greater;
fewer intermolecular bonds/forces to break / vaporization is endothermic / gaseous fuel has greater enthalpy than liquid fuel / *OWTTE*; **[2]**
*M2 cannot be scored if M1 is incorrect.*

Q# 12/ IB Chem/2011/s/TZ1/Paper 2 Section A/Standard Level/

**1.** (a) *amount of energy required to break bonds of reactants*
$3 \times 413 + 358 + 464 + 1.5 \times 498$ (kJ mol$^{-1}$) / 2808 (kJ mol$^{-1}$);
*amount of energy released during bond formation of products*
$4 \times 464 + 2 \times 746$ (kJ mol$^{-1}$) / 3348 (kJ mol$^{-1}$);
$\Delta H = -540$ (kJ mol$^{-1}$); **[3]**
*Award **[3]** for correct final answer.*
*Award **[2]** for (+)540.*
*If old Data Booklet is used accept answer: –535 (kJ mol$^{-1}$) or award **[2]** for (+)535.*

(b) (i) $m$(methanol) = (80.557 – 80.034) = 0.523 (g);
$n(\text{methanol}) = \left(\frac{0.523\text{ g}}{32.05\text{ g mol}^{-1}}\right) = 0.0163$ (mol); **[2]**
*Award **[2]** for correct final answer.*




(ii) $\Delta T = (26.4 - 21.5) = 4.9$ (K);
$q = (mc\Delta T =)\ 20.000 \times 4.18 \times 4.9$ (J) / $20.000 \times 4.18 \times 4.9 \times 10^{-3}$ (kJ);
0.41 (kJ); **[3]**
*Award **[3]** for correct final answer.*

(iii) $\Delta H_c^{\ominus} = -\frac{0.41\text{ (kJ)}}{0.0163\text{ (mol)}} / -25153\text{ (J mol}^{-1}\text{)}$;
$= -25$ (kJ mol$^{-1}$); **[2]**
*Award **[2]** for correct final answer.*
*Award **[1]** for (+)25 (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 gas in the reaction; **[1]**

(ii) not all heat produced transferred to water / heat lost to surroundings/environment / *OWTTE* / incomplete combustion (of methanol) / water forms as $H_2O(l)$ instead of $H_2O(g)$; **[1]**
*Do not allow just "heat lost".*

Q# 13/ IB Chem/2010/w/tz0/Paper 2 Section B/Standard Level/q4

(d) bonds broken: 4 N–H, N–N, O=O / +2220 (kJ mol$^{-1}$);
bonds formed: N≡N, 4 O–H / –2801 (kJ mol$^{-1}$);
–581 (kJ mol$^{-1}$); **[3]**
*Award **[3]** for correct final answer.*

Q# 14/ IB Chem/2010/w/TZ0/Paper 2 Section A/Standard Level/

**1.** (a) all heat is transferred to water/copper sulfate solution / no heat loss;
specific heat capacity of zinc is zero/negligible / no heat is absorbed by the zinc;
density of water/solution = 1.0 / density of solution = density of water;
heat capacity of cup is zero / no heat is absorbed by the cup;
specific heat capacity of solution = specific heat capacity of water;
temperature uniform throughout solution; **[2 max]**
*Award **[1]** each for any two.*
*Accept energy instead of heat.*

(b) (i) $T_{\text{final}} = 73.0\,(^{\circ}C)$;
*Allow in the range 72 to 74 (°C).*
$\Delta T = 48.2\,(^{\circ}C)$; **[2]**
*Allow in the range 47 to 49 (°C).*
*Award **[2]** for correct final answer*
*Allow ECF if $T_{final}$ or $T_{initial}$ correct.*

(ii) temperature decreases at uniform rate (when above room temperature) / *OWTTE*; **[1]**

(iii) 10.1 (kJ); **[1]**
*Allow in the range 9.9 to 10.2 (kJ).*

(c) $\left(n_{Zn} = n_{CuSO_4} = \frac{1.00 \times 50.0}{1000}\right) = 0.0500$ (mol); **[1]**

(d) –201 (kJ mol$^{-1}$); **[1]**
*Allow in the range –197 to – 206 (kJ mol$^{-1}$).*
*Value must be negative to award mark.*



(e) (i) the more reactive the metal the more negative the enthalpy change/the more exothermic the reaction / *OWTTE*; *[1]*
*Do not accept greater enthalpy change.*

(ii) any curve with positive gradient which passes through $\Delta H = 0$ at Cu; *[1]*
*Allow point graph or histogram.*
*Accept either positive or zero enthalpy change for Ag.*

Q# 15/ IB Chem/2010/s/tz1/Paper 2 Section B/Standard Level/

**6.** (a) (i) $100 \times 4.18 \times 35.0$;
14630 J / 14600 J / 14.6 kJ; *[2]*
*Award [2] for correct final answer.*
*No ECF here if incorrect mass used.*

(ii) $\frac{1.78}{46.08} = 0.0386$ mol;

$\frac{14.6}{0.0386} = (-)378$ kJ mol$^{-1}$; *[2]*

*Accept (–)377 and (–)379 kJ mol$^{-1}$.*
*Award [2] for correct final answer.*

(iii) heat loss;
incomplete combustion;
heat absorbed by calorimeter not included; *[2 max]*
*Accept other sensible suggestions.*

Q# 16/ IB Chem/2009/w/TZ0/Paper 2 Section A/Standard Level/

**3.** $(C_2H_6(g) + 3\frac{1}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l))$ $\Delta H^\ominus = -1560$;
$(H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g))$ $\Delta H^\ominus = +286$;
$(2CO_2(g) + 2H_2O(l) \rightarrow C_2H_4(g) + 3O_2(g))$ $\Delta H^\ominus = +1411$;
$(C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g))$ $\Delta H^\ominus = +137$ (kJ); *[4]*
*Allow other correct methods.*
*Award [2] for –137.*
*Allow ECF for the final marking point.*

Q# 17/ IB Chem/2009/s/TZ1/Paper 2 Section A/Standard Level/

**2.** (a) energy required = C=C + H–H/612 + 436 and
energy released = C–C + 2(C–H)/347 + 2(413) /
energy required = C=C + H–H + 4(C–H)/612 + 436 + 4(413) and
energy released = C–C + 6(C–H)/347 + 6(413);

$\Delta H = (1048 - 1173)/(2700 - 2825) = -125$ kJ mol$^{-1}$; *[2]*

(b) $\Delta H = -1411 + (-286) - (-1560) = -137$ kJ mol$^{-1}$; *[1]*

(c) the actual values for the specific bonds may be different to the average values / the combustion values referred to the specific compounds / *OWTTE*; *[1]*

(d) (i) –125 kJ mol$^{-1}$; *[1]*

(ii) average bond enthalpies do not apply to the liquid state / *OWTTE*;
the enthalpy of vaporization/condensation of cyclohexene and cyclohexane / *OWTTE*; *[2]*




Q# 18/ IB Chem/2008/w/tz0/Paper 2 Section B/Standard Level/

**4.** (a) (i) energy change to break/make (one mole) of bonds (in molecule) in gaseous state; averaged over similar compounds; *[2]*

(ii) H–H bond is only present in one molecule/*OWTTE*; *[1]*

(iii) bonds broken: C=C and $\frac{1}{2}$ O=O / 612 + 248 / 860;
bonds formed: C–C and 2 C–O / 348 + 720 / 1068;
$\Delta H (= 860 - 1068) = -208$ kJ/kJ mol$^{-1}$; *[3]*
*Correct final answer award [3] marks.*
*208 kJ/kJ mol$^{-1}$ or +208 kJ/kJ mol$^{-1}$ award [2] marks.*
*ECF from bonds broken and formed.*
*No penalty for including 4 C–H bonds in both bonds broken and formed.*

(iv) *diagram showing*

vertical axis labelled enthalpy/energy;
*Do not accept enthalpy change.*

reactants line above products line and $\Delta H$; *[2]*
*ECF from sign in (iii).*

*Ignore connecting lines, intermediates, activated complexes.*
*Enthalpy change value not needed, but if in wrong place, then deduct [1] mark.*

(v) negative;
decrease in gas moles/molecules/volume/decrease in disorder/increase in order; *[2]*

(b) (i) *for first equation* $8 \times -394 / -3152$;
*for second equation* $9 \times -286 / -2574$;
*for third equation* +5512;
$\Delta H^\ominus = -214$ (kJ or kJ mol$^{-1}$); *[4]*
*No penalty for missing units, but penalize wrong units by –1(U).*
*Correct final answer award [4] marks.*
*All marks can be scored from enthalpy cycle.*
*Consider ECF for final mark.*

(ii) II;
heat evolved during condensation / when hydrogen bonds/intermolecular forces form / latent heat given out / *OWTTE*; *[2]*

Q# 19/ IB Chem/2008/s/TZ1/Paper 2 Section A/Standard Level/

**3.** (a) $(2C(s) + 2O_2(g) \rightarrow 2CO_2(g))$ $\Delta H^\ominus = -788$ kJ;
$(2H_2(g) + O_2(g) \rightarrow 2H_2O(l))$ $\Delta H^\ominus = -572$ kJ;
$(2CO_2(g) + 2H_2O(l) \rightarrow CH_3COOH(l) + 2O_2(g))$ $\Delta H^\ominus = +876$ kJ;

$2C(s) + 2H_2(g) + O_2(g) \rightarrow CH_3COOH(l)$ $\Delta H^\ominus = -484$ (kJ); *[4]*
*Award [4] for correct final answer.*

Q# 20/ IB Chem/2007/w/tz0/Paper 2 Section B/Standard Level/q5a

(ii) products more stable than reactants / reactants less stable than products;
products lower in energy / reactants higher in energy; **[2]**

(iii) (overall) bonds in reactants weaker / (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) / *OWTTE*;

**OR**

bond breaking is endothermic / requires energy and bond making is exothermic / releases energy;
stronger bonds in product mean process is exothermic overall; **[2]**

Q# 21/ IB Chem/2007/s/TZ1/Paper 2 Section A/Standard Level/

**3.** (a) amount of energy needed to break one mole of (covalent) bonds;
in the gaseous state;
average calculated from a range of compounds; **[2 max]**
*Award [1] each for any two points above.*

(b) bonds broken: $612+2\times348+8\times412+6\times496$ / 7580 kJ $mol^{-1}$;
bonds made: $8\times743+8\times463$ / 9648 kJ $mol^{-1}$;
(bonds broken – bonds made =) $\Delta H = -2068$ (kJ $mol^{-1}$); **[3]**
*Award [3] for the correct answer.*
*Allow full ECF – 1 mistake equals 1 penalty.*
*Allow kJ but not other wrong units.*

(c) same/equal, because the same bonds are being broken and formed; **[1]**

(e) products more stable than reactants;
bonds are stronger in products than reactants / $H_P < H_R$ / enthalpy / stored energy of products less than reactants; **[2]**

Q# 22/ IB Chem/2007/s/TZ1/Paper 2 Section A/Standard Level/

**3.** (a) amount of energy needed to break one mole of (covalent) bonds;
in the gaseous state;
average calculated from a range of compounds; **[2 max]**
*Award [1] each for any two points above.*

(b) bonds broken: $612+2\times348+8\times412+6\times496$ / 7580 kJ $mol^{-1}$;
bonds made: $8\times743+8\times463$ / 9648 kJ $mol^{-1}$;
(bonds broken – bonds made =) $\Delta H = -2068$ (kJ $mol^{-1}$); **[3]**
*Award [3] for the correct answer.*
*Allow full ECF – 1 mistake equals 1 penalty.*
*Allow kJ but not other wrong units.*

(c) same/equal, because the same bonds are being broken and formed; **[1]**

(e) products more stable than reactants;
bonds are stronger in products than reactants / $H_P < H_R$ / enthalpy / stored energy of products less than reactants; **[2]**




Q# 23/ IB Chem/2006/w/TZ0/Paper 2 Section A/Standard Level/

**2.** (a) energy needed to break (1 mol of) a bond in a gaseous molecule;
averaged over similar compounds; **[2]**

(b) bonds broken identified as C–O and N–H;
bonds formed identified as C–N and O–H;
$\Delta H = 748 - 768$ (kJ);
$= -20$ kJ / kJ $mol^{-1}$ *(units needed for this mark)*; **[4]**
*If wrong bonds identified apply ECF to 3rd and 4th marks.*
*Accept answer based on breaking and making all bonds.*
*Award [4] for correct final answer.*
*Award max [3] if only one bond missed.*
*Answer of 20 or +20 kJ ($mol^{-1}$) scores [3].*

Q# 24/ IB Chem/2006/s/TZ1/Paper 2 Section A/Standard Level/

**1.** (a) (Amount of energy required to break bonds of reactants)
$8\times412+2\times348+612+6\times496$ / 7580 (kJ $mol^{-1}$);

(Amount of energy released during bond formation)
$4\times2\times743+4\times2\times463$ / 9648 (kJ $mol^{-1}$);

$\Delta H = -2068$ (kJ or kJ $mol^{-1}$); **[3]**
*ECF from above answers.*
*Correct answer scores [3].*
*Award [2] for (+)2068.*
*If any other units apply –1(U), but only once per paper.*

(b) exothermic **and** $\Delta H^{\ominus}$ is negative / energy is released; **[1]**
*Apply ECF to sign of answer in part (a).*
*Do not mark if no answer to (a).*

(c) $-1\times\Delta H_1$ / 676;
$1\times\Delta H_2$ / –394;
$2\times\Delta H_3$ / – 484;

$\Delta H_4 = -202$ (kJ $mol^{-1}$); **[4]**
*Accept alternative methods.*
*Correct answers score [4].*
*Award [3] for (+)202 or (+)40 (kJ/kJ $mol^{-1}$).*
*–1(U) if units incorrect (ignore if absent).*

Q# 25/ IB Chem/2005/w/tz0/Paper 2 Section B/Standard Level/

**5.** (a) (i) standard enthalpy (change) of reaction;
(temperature) increase;
reaction is exothermic/sign of $\Delta H^{\ominus}$ is negative; **[3]**

(ii) more (negative);
heat given out when gas changes to solid/solid has less enthalpy than gas/*OWTTE*; **[2]**

(iii) –389 kJ; **[1]**

(iv) negative;
fewer gas particles/decrease in gas moles;
decrease in entropy/increase in order/decrease in disorder; **[3]**

(b) (i) the energy needed to break one bond;
(in a molecule in the) gaseous state;
value averaged using those from similar compounds; **[3]**

(ii) it is an element/no other species with just a Br-Br bond/*OWTTE*; *[1]*

(iii) (sum bonds broken =) 412 + 193 = 605;
(sum bonds formed =) 276 + 366 = 642;
($\Delta H^{\ominus}$ =) – 37 kJ; *[3]*
*Award [3] for correct final answer.*
*Award [2] for "+ 37".*
*Accept answer based on breaking and making extra C-H bonds.*

(iv)

Enthalpy | $CH_4 + Br_2$ / $CH_3Br + HBr$ ; *[2]*

*Award [1] for enthalpy label and two horizontal lines, [1] for reactants higher than products.*
*ECF from sign in (iii), ignore any higher energy level involving atoms.*

(v) (about) the same/similar;
same (number and type of) bonds being broken and formed; *[2]*

Q# 26/ IB Chem/2005/s/TZ1/Paper 2 Section A/Standard Level/

**1.** (a) exothermic because temperature rises / heat is released; *[1]*

(b) to make any heat loss as small as possible / so that all the heat will be given out very quickly; *[1]*
*Do not accept "to produce a faster reaction".*

(c) heat released = mass × specific heat capacity × temp increase / $q = mc\Delta T$ / 100 × 4.18 × 3.5;
= 1463 J / 1.463 kJ; *(allow 1.47 kJ if specific heat = 4.2)*
amount of KOH / HCl used = 0.500 × 0.050 = 0.025 mol;
$\Delta H = (1.463 \div 0.025) = -58.5$ (kJ mol$^{-1}$); *(minus sign needed for mark)* *[4]*
*Use ECF for values of q and amount used.*
*Award [4] for correct final answer.*
*Final answer of 58.5 or +58.5 scores [3].*
*Accept 2,3 or 4 significant figures.*

(d) heat loss (to the surroundings);
insulate the reaction vessel / use a lid / draw a temperature versus time graph; *[2]*

(e) 3.5 °C / temperature change would be the same;
amount of base reacted would be the same / excess acid would not react / KOH is the limiting reagent; *[2]*

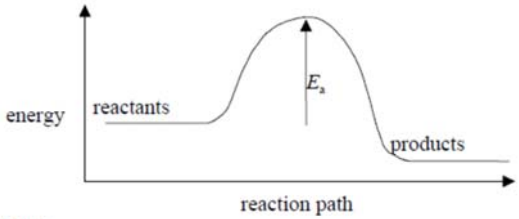
Q# 27/ IB Chem/2004/w/TZ0/Paper 2 Section A/Standard Level/

**3.** (a) energy for the conversion of a gaseous molecule into (gaseous) atoms;
(average values) obtained from a number of similar bonds / compounds / *OWTTE*;
$CH_4(g) \rightarrow C(g) + 4H(g)$; *[3]*
*State symbols needed.*

(b) (bond breaking) = 1890 / 654;
(bond formation) = 2005 / 769;
enthalpy = –115 (kJ mol$^{-1}$); *[3]*
*Allow ECF from bond breaking and forming.*
*Award [3] for correct final answer.*
*Penalize [1] for correct answer with wrong sign.*



(d)

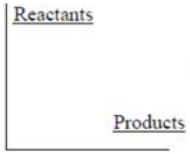



exothermic shown;
activation energy / $E_a$ shown; *[2]*
*Allow ECF from (b).*

Q# 28/ IB Chem/2004/s/tz1/Paper 2 Section B/Standard Level/

**5.** (a) (i) $\Delta H = \Delta H_f$ (products) – $\Delta H_f$ (reactants) / = (–1669) – (–822)
= –847 kJ
*Ignore units;*
exothermic (ECF from sign of $\Delta H$); *[3]*

(ii)




*ignore lines connecting reactants and products*

*[1] for the diagram*
*ECF from sign of ΔH in (a) (i)*
298 K / 25 °C AND 1 atm / 101(.3) kPa; *[2]*
*Both needed for the mark.*

Q# 29/ IB Chem/2003/w/TZ0/Paper 2 Section A/Standard Level/

**1.** (a) $C(s) + 2F_2(g) \rightarrow CF_4(g)$ $\Delta H_1 = -680$ kJ;
$4F(g) \rightarrow 2F_2(g)$ $\Delta H_2 = 2(-158)$ kJ;
$C(g) \rightarrow C(s)$ $\Delta H_3 = -715$ kJ;
*Accept reverse equations with* $+\Delta H$ *values.*

$C(g) + 4F(g) \rightarrow CF_4(g)$ $\Delta H = -1711$ kJ, so average bond enthalpy $= \frac{-1711}{4}$ *[4]*
$= -428$ kJ mol$^{-1}$;
*Accept + or – sign.*

*Lots of ways to do this! The correct answer is very different from the value in the Data Booklet, so award [4] for final answer with/without sign units not needed, but deduct [1] if incorrect units. Accept answer in range of 427 to 428 without penalty for sig figs.*

*If final answer is not correct use following;*
*Award [1] for evidence of cycle or enthalpy diagram or adding of equations.*
*Award [1] for or* $2F_2(g) \rightarrow 4F(g)$ *2 x 158 seen.*
*Award [1] for dividing 1711 or other value by 4.*

Q# 30/ IB Chem/2001/s/TZ1/Paper 2 Section A/Standard Level/

**1.** (a) $C + \frac{1}{2}O_2 \rightarrow CO$ *(ignore state symbols)* ***[1]***

some evidence of working *e.g.* cycle or changing sign of $\Delta H$ ***[1]***

$-110.5$ *(units not required)* ***[1]***

*(−110.5 on its own scores [3])* ***[3 max]***

(b) absorbs heat / $\Delta H$ is positive / absorbs energy / products have more energy than reactants. ***[1]***

(c) (i) Breaking bonds $C \equiv C$; $4(C—H)$; $3(O = O)$ ***[1]***

Making bonds $2(O = C = O)$; $2(H—O—H)$ ***[1]***

Breaking +3748 Making −4824 ***[1]***

Enthalpy of combustion = − 1076 *(+1076 scores [3 max])* ***[1]*** ***[4 max]***

*(In the absence of any credit, award **[1]** for breaking (+) **and** making (−) or $\Delta H_c = H_{products} - H_{reactants}$.)*

(ii) Exothermic since $\Delta H_c$ is negative (*NB consequential on sign in (c) (ii)*). ***[1]***

*(If (c) (i) is not attempted, allow exothermic because hydrocarbon combustion gives out heat / OWTTE).*

Q# 31/ IB Chem/2000/s/tz1/Paper 2 Section B/Standard Level/

**5.** (a) (i) Enthalpy of products – enthalpy of reactants ***[1]***
Specified temperature (*e.g.* 298 K) ***[1]***
Specified pressure (*e.g.* $1.013 \times 10^5$ Pa) ***[1]***

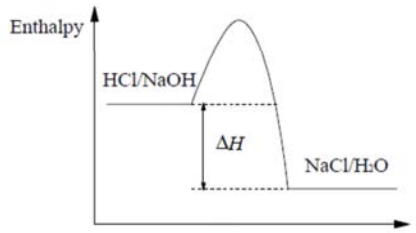
(ii) Take a **known volume** of sodium hydroxide solution ***[1]***
of **known molarity/concentration** ***[1]***
**record its temperature** ***[1]***
and place in an **insulated vessel**. ***[1]***
Add a solution of HCl, of known temperature, such that an **equal/excess number of moles is added**. ***[1]***
**Stir/mix**, ***[1]***
record the maximum **rise in temperature**. ***[1]***

Need: Temperature rise ***[1]***.
$\Delta H$ = Total mass (or volume) × temperature rise × specific heat capacity ***[1]***

Divide by number of moles (of limiting reactant if excess or of other reactant used) ***[1]*** (*i.e.* factor to convert to 1 mole).

*(Any 9 points)*

(iii)

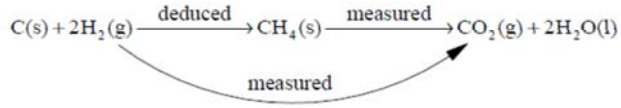





Overall diagram: enthalpy label and two different levels; ***[1]***
labelled enthalpy levels – NaCl / $H_2O$ must be lower; ***[1]***
enthalpy change shown; ***[1]***
products more stable than reactants/reaction exothermic. ***[1]***

(b) Hess's Law enables enthalpy changes which cannot be found experimentally to be calculated from other experimental results. ***[1]***
Suitable example; ***[1]***
Intermediate stage shown; ***[1]***
Show which enthalpies can be measured, hence which deduced. ***[1]***

Example:




*(Award 1st mark for overall equation, i.e. example; award 2nd mark for intermediate; award 3rd mark for showing $\Delta H_f$ deduced from enthalpies of combustion.)*

Q# 32/ IB Chem/1999/w/TZ0/Paper 2 Section A/Standard Level/

**1.** (a) $CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$

**OR**

$Cu^2 + Zn \rightarrow Zn^{2+} + Cu$ ***[1]***

***States not necessary to gain mark***

(b) $\text{Amount of Zn} = \frac{1.20}{65.37} = 0.018 \text{ moles}$

$\text{Amount of } Cu^{2+} = \frac{50}{1000} \times 0.200 = 0.010 \text{ moles}$ ***[1]***

**therefore Zn is in excess** ***[1]***

(c) At point A the heat being given out by **the reaction is equal to the heat being lost to the surroundings.** ***[1]***

*Do **not** give the mark for "the reaction is finished".*

(d) Correct extrapolation to when the zinc was added. ***[1]***

*Give no marks if the line is extrapolated to the Y axis.*

**Temperature rise = 26.7 − 17.0 = 9.7° C** ***[1]***

*Accept 26.7 ± 0.1° C giving 9.6 to 9.8° C*

(e) Heat = 9.7 × 4.18 × 50 *Give credit if 51.2 g taken as mass* ***[1]***
= 2027.3 J = 2030 J *Answer must be given to 3 s.f. to gain mark* ***[1]***

*Consequential markings from (d)*

(f) $\Delta H = -2030 \times 100$ J
$= -203$ kJ mol$^{-1}$ *Must have – sign* *[1]*

(g) $\text{Error} = \frac{218 - 203}{218} \times 100 = 6.9\ \%$ *[1]*

(h) ***[1 mark]*** *for any valid reason*

*e.g.* solution assumed to have same specific heat capacity as 50 g of water;
Heating of metal (Cu + excess Zn), thermometer *etc.* ignored.
(not carried out under standard conditions (not on SL syllabus);) *[1]*

Q# 33/ IB Chem/1999/s/TZ1/Paper 2 Section A/Standard Level/

**2.** (a) **energy** needed to **break** ***[1 mark]***

(a mole of covalent bonds)
in the **gaseous** state ***[1 mark]***

(b) −698 (kJ mol$^{-1}$) ***value [1 mark]***
***sign [1 mark]***

+698 scores ***[1 mark]***