**IB Chemistry SL**

**Topic 5 Questions**

**1.** What energy changes occur when chemical bonds are formed and broken?

A. Energy is absorbed when bonds are formed and when they are broken.

B. Energy is released when bonds are formed and when they are broken.

C. Energy is absorbed when bonds are formed and released when they are broken.

D. Energy is released when bonds are formed and absorbed when they are broken.

**2.** 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 of Al = 0.90 J g–1K–1)

A. 0.36

B. 2.3

C. 9.0

D. 11

**3.** Using the equations below:

C(s) + O2(g) → CO2(g) *∆H* = –390 kJ  
Mn(s) + O2(g) → MnO2(s) *∆H* = –520 kJ

what is ∆*H* (in kJ) for the following reaction?

MnO2(s) + C(s) → Mn(s) + CO2(g)

A. 910

B. 130

C. –130

D. –910

**4.** Which statements about exothermic reactions are correct?

I. They have negative Δ*H* values.

II. The products have a lower enthalpy than the reactants.

III. The products are more energetically stable than the reactants.

A. I and II only

B. I and III only

C. II and III only

D. I, II and III

**5.** A sample of a metal is heated. Which of the following are needed to calculate the heat absorbed by the sample?

I. The mass of the sample

II. The density of the sample

III. The specific heat capacity of the sample

A. I and II only

B. I and III only

C. II and III only

D. I, II and III

**6.** The average bond enthalpies for O—O and O==O are 146 and 496 kJ mol–1 respectively.   
What is the enthalpy change, in kJ, for the reaction below?

H—O—O—H(g) → H—O—H(g) + ½O==O(g)

A. – 102 B. + 102

C. + 350 D. + 394

**7.** When the solids Ba(OH)2 and NH4SCN are mixed, a solution is produced and the temperature drops.

Ba(OH)2 (s) + 2 NH4SCN (s) → Ba(SCN)2 (aq) + 2 NH3 (g) + 2 H2O (l)

Which statement about the energetics of this reaction is correct?

A. The reaction is endothermic and Δ*H* is negative.

B. The reaction is endothermic and Δ*H* is positive.

C. The reaction is exothermic and Δ*H* is negative.

D. The reaction is exothermic and Δ*H* is positive.

**8.** Using the equations below

Cu (s) + O2 (g) → CuO (s) ∆*H*~~ο~~= –156 kJ

.2 Cu (s) + O2 (g) → Cu2O (s) ∆*H*~~ο~~= –170 kJ

what is the value of ∆*H*~~ο~~ (in kJ) for the following reaction?

2 CuO(s) → Cu2O(s) + O2 (g)

A. 142 B. 15

C. –15 D. –142

**9.** Which of the quantities in the enthalpy level diagram below is (are) affected by the use of a catalyst?

  
Time

A. I only

B. III only

C. I and II only

D. II and III only

**10.** Consider the following equations.

Mg(s) + O2(g) → MgO(s) ∆*H*~~ο~~ = –602 kJ  
H2(g) + O2(g) → H2O(g) ∆*H*~~ο~~ = –242 kJ

What is the ∆*H*° value (in kJ) for the following reaction?

MgO(s) + H2(g) → Mg(s) + H2O(g)

A. –844 B. –360

C. +360 D. +844

**11.** For which of the following is the sign of the enthalpy change different from the other three?

A. CaCO3(s) → CaO(s) + CO2(g)

B. Na(g) → Na+(g) + e–

C. CO2(s) → CO2(g)

D. 2Cl(g) → Cl2(g)

**12.** Separate solutions of HCl(aq) and H2SO4(aq) of the same concentration and same volume were completely neutralized by NaOH(aq). *X* kJ and *Y* kJ of heat were evolved respectively.   
Which statement is correct?

A. *X* = *Y* B. *Y* = 2*X*

C. *X* = 2*Y* D. *Y* = 3*X*

**13.** Which statements are correct for an endothermic reaction?

I. The system absorbs heat.

II. The enthalpy change is positive.

III. The bond enthalpy total for the reactants is greater than for the products.

A. I and II only

B. I and III only

C. II and III only

D. I, II and III

**14.** The mass *m* (in g) of a substance of specific heat capacity *c* (in J g–1 K–1 ) increases by *t°*C. What is the heat change in J?

A. *mct*

B. *mc*(*t* + 273)

C. 

D. 

**15.** The average bond enthalpy for the C―H bond is 412 kJ mol–1. Which process has an enthalpy change closest to this value?

A. CH4(g) → C(s) + 2H2(g)

B. CH4(g) → C(g) + 2H2(g)

C. CH4(g) → C(s) + 4H(g)

D. CH4(g) → CH3(g) + H(g)

**16.** The following equation shows the formation of magnesium oxide from magnesium metal.

2Mg(s) + O2(g) → 2MgO(s) Δ*HӨ* = –1204kJ

Which statement is correct for this reaction?

A. 1204 kJ of energy are released for every mol of magnesium reacted.

B. 602 kJ of energy are absorbed for every mol of magnesium oxide formed.

C. 602 kJ of energy are released for every mol of oxygen gas reacted.

D. 1204 kJ of energy are released for every two mol of magnesium oxide formed.

**17.** The following equations show the oxidation of carbon and carbon monoxide to carbon dioxide.

C(s) + O2(g) → CO2(g) Δ*H*Ө = –*x* kJ mol–1

CO(g) + O2(g) → CO2(g) Δ*H*Ө = –*y* kJ mol–l

What is the enthalpy change, in kJ mol–1, for the oxidation of carbon to carbon monoxide?

C(s) + O2(g) → CO(g)

A. *x* + *y* B. – *x* – *y*

C. *y* – *x* D. *x* – *y*

**18.** A simple calorimeter was used to determine the enthalpy of combustion of ethanol. The experimental value obtained was –920 kJ mol–1. The Data Booklet value is –1371 kJ mol–1. Which of the following best explains the difference between the two values?

A. incomplete combustion of the fuel

B. heat loss to the surroundings

C. poor ventilation in the laboratory

D. inaccurate temperature measurements

**19.** For the reaction

2H2(g) + O2(g) → 2H2O(g)

the bond enthalpies (in kJ mol–1) are

|  |  |
| --- | --- |
| H–H | *x* |
| O=O | *y* |
| O–H | *z* |

Which calculation will give the value, in kJ mol–1, of Δ*H*Ө for the reaction?

A. 2*x* + *y* –2*z*

B. 4*z* – 2*x* – *y*

C. 2*x* + *y* – 4*z*

D. 2*z* –2*x* – *y*

**20.** Which statement about bond enthalpies is correct?

A. Bond enthalpies have positive values for strong bonds and negative values for weak bonds.

B. Bond enthalpy values are greater for ionic bonds than for covalent bonds.

C. Bond breaking is endothermic and bond making is exothermic.

D. The carbon–carbon bond enthalpy values are the same in ethane and ethene.

**21.** An equation for a reaction in which hydrogen is formed is

CH4 + H2O → 3H2 + CO Δ*H*Ө *=* +210 kJ

Which energy change occurs when 1 mol of hydrogen is formed in this reaction?

A. 70 kJ of energy are absorbed from the surroundings.

B. 70 kJ of energy are released to the surroundings.

C. 210 kJ of energy are absorbed from the surroundings.

D. 210 kJ of energy are released to the surroundings.

**22.** The equations and enthalpy changes for two reactions used in the manufacture of sulfuric acid are:

S(s) O2(g) → SO2(g) Δ*H*Ө = –300 kJ

2SO2(g) + O2(g) → 2SO3(g) Δ*H*Ө = –200 kJ

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

2S(s) + 3O2(g) → 2SO3(g)

A. –100

B. –400

C. –500

D. –800

**23.** Approximate values of the average bond enthalpies, in kJ mol–1, of three substances are:

|  |  |
| --- | --- |
| H–H | 430 |
| F–F | 155 |
| H–F | 565 |

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

2HF → H2 + F2

A. +545

B. +20

C. –20

D. –545

**24.** The standard enthalpy change of formation values of two oxides of phosphorus are:

P4(s) + 3O2(g) → P4O6(s) Δ*H*Өf= –1600 kJ mol–1

P4(s) + 5O2(g) → P4O10(s) Δ*H*Өf= –3000 kJ mol–1

What is the enthalpy change, in kJ mol–1, for the reaction below?

P4O6(s) + 2O2(g) → P4O10(s)

A. +4600

B. +1400

C. –1400

D. –4600

**25.** Which statement is correct for an endothermic reaction?

A. The products are more stable than the reactants and Δ*H* is positive.

B. The products are less stable than the reactants and Δ*H* is negative.

C. The reactants are more stable than the products and Δ*H* is positive.

D. The reactants are less stable than the products and Δ*H* is negative.

**26.** Which statement is correct about the reaction shown?

2SO2(g) + O2(g) → 2SO3(g) Δ*H* = –196 kJ

A. 196 kJ of energy are released for every mole of SO2(g) reacted.

B. 196 kJ of energy are absorbed for every mole of SO2(g) reacted.

C. 98 kJ of energy are released for every mole of SO2(g) reacted.

D. 98 kJ of energy are absorbed for every mole of SO2(g) reacted.

**27.** Which statements are correct for all exothermic reactions?

I. The enthalpy of the products is less than the enthalpy of the reactants.

II. The sign of Δ*H* is negative.

III. The reaction is rapid at room temperature.

A. I and II only

B. I and III only

C. II and III only

D. I, II and III

**28.** Consider the specific heat capacity of the following metals.

|  |  |
| --- | --- |
| Metal | Specific heat capacity / J kg–1 K–1 |
| Cu | 385 |
| Ag | 234 |
| Au | 130 |
| Pt | 134 |

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

B. Ag

C. Au

D. Pt

**29.** Consider the following reactions.

S(s) + O2(g) → SO3(g) Δ*H*Ө = -395 kJ mol-1

SO2(s) + O2(g) → SO3(g) Δ*H*Ө = -98 kJ mol-1

What is the Δ*H*Ө value (in kJ mol–1) for the following reaction?

S(s) + O2(g) → SO2 (g)

A. –297

B. +297

C. – 493

D. +493

**30.** Which statement is correct for an endothermic reaction?

A. Bonds in the products are stronger than the bonds in the reactants.

B. Bonds in the reactants are stronger than the bonds in the products.

C. The enthalpy of the products is less than that of the reactants.

D. The reaction is spontaneous at low temperatures but becomes non-spontaneous at high temperatures.

**31.** According to the enthalpy level diagram below, what is the sign for Δ*H* and what term is used to refer to the reaction?



|  |  |  |
| --- | --- | --- |
|  | Δ*H* | reaction |
| A. | positive | endothermic |
| B. | negative | exothermic |
| C. | positive | exothermic |
| D. | negative | endothermic |

**32.** When 40 joules of heat are added to a sample of solid H2O at –16.0°C the temperature increases to –8.0°C. What is the mass of the solid H2O sample?

[Specific heat capacity of H2O(s) = 2.0 J g–1 K–1]

A. 2.5 g

B. 5.0 g

C. 10 g

D. 160 g

**33.** The Δ*H*Ө values for the formation of two oxides of nitrogen are given below.

N2(g) + O2(g) → NO2(g) Δ*H*Ө = –57 kJ mol–1

N2(g) + 2O2(g) → N2O4(g) Δ*H*Ө = +9 kJ mol–1

Use these values to calculate Δ*H*Ө for the following reaction (in kJ):

2NO2(g) → N2O4(g)

A. –105

B. – 48

C. +66

D. +123

**34.** How much energy, in joules, is required to increase the temperature of 2.0 g of aluminium from 25 to 30°C? (Specific heat of Al = 0.90 J g–1 K–1).

A. 0.36

B. 4.5

C. 9.0

D. 54

**35.** Which combination is correct for a chemical reaction that absorbs heat from the surroundings?

|  |  |  |
| --- | --- | --- |
|  | **Type of reaction** | **Δ*H* at constant pressure** |
| A. | Exothermic | Positive |
| B. | Exothermic | Negative |
| C. | Endothermic | Positive |
| D. | Endothermic | Negative |

**36.** Using the equations below:

C(s) + O2(g) → CO2(g) ∆*H*~~ο~~ = –394 kJ mol–1

Mn(s) + O2(g) → MnO2(s) ∆*H*~~ο~~ = –520 kJ mol–1

What is ∆*H*, in kJ, for the following reaction?

MnO2(s) + C(s) → Mn(s) + CO2(g)

A. 914

B. 126

C. –126

D. –914

**37.**



The diagram shows the distribution of energy for the molecules in a sample of gas at a given temperature, T1.

(a) In the diagram *E*a represents the *activation energy* for a reaction. Define this term.

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(1)

(b) On the diagram above draw another curve to show the energy distribution for the same gas at a higher temperature. Label the curve T2.

(2)

(c) With reference to your diagram, state and explain what happens to the rate of a reaction when the temperature is increased.

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(2)

(Total 5 marks)

**38.** (a) Define the term *average bond enthalpy*, illustrating your answer with an equation for methane, CH4.

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(3)

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

CH4(g) + Cl2(g) → CH3Cl(g) + HCl(g)

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

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(3)

(c) Explain why no reaction takes place between methane and chlorine at room   
temperature unless the reactants are sparked, exposed to UV light or heated.

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(2)

(d) Draw an enthalpy level diagram for this reaction.

(2)

(Total 10 marks)

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

KOH(aq) + HCl(aq) → KCl(aq)+ H2O(l)

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

50.0 cm3 of a 0.500 mol dm–3 solution of KOH was mixed rapidly in a glass beaker with 50.0 cm3 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.

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(1)

(b) Explain why the solutions were mixed rapidly.

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(1)

(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.

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

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(2)

(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.

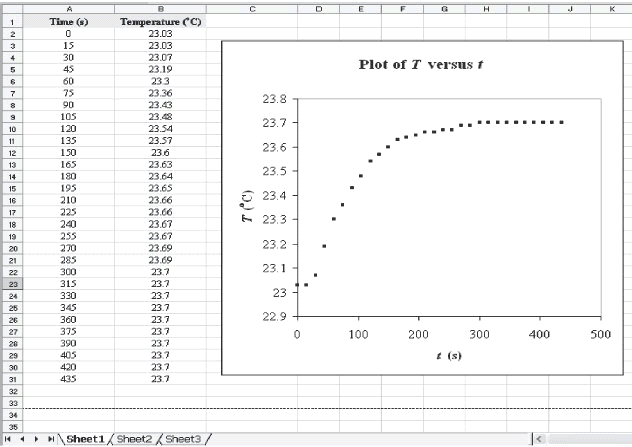
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(2)

(Total 10 marks)

**40.** The data below is from an experiment used to measure the enthalpy change for the combustion of 1 mole of sucrose (common table sugar), C12H22O11(s). The time-temperature data was taken from a data-logging software programme.



Mass of sample of sucrose, *m* = 0.4385 g

Heat capacity of the system, *C*system = 10.114 kJ K–1

(a) Calculate Δ*T*, for the water, surrounding the chamber in the calorimeter.

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(1)

(b) Determine the amount, in moles, of sucrose.

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(1)

(c) (i) Calculate the enthalpy change for the combustion of 1 mole of sucrose.

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(1)

(ii) Using Table 12 of the Data Booklet, calculate the percentage experimental error based on the data used in this experiment.

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(1)

(d) A hypothesis is suggested that TNT, 2-methyl-1,3,5-trinitrobenzene, is a powerful explosive because it has:

• a large enthalpy of combustion

• a high reaction rate

• a large volume of gas generated upon combustion

Use your answer in part (c)(i) and the following data to evaluate this hypothesis:

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Equation for combustion** | **Relative rate of combustion** | **Enthalpy of combustion / kJ mol–1** |
| Sucrose | C12H22O11 (s) + 12 O2 (g) → 12 CO2 (g) + 11 H2O (g) | Low |  |
| TNT | 2 C7H5N3O6 (s) → 7 CO(g) + 7 C (s) + 5 H2O (g) + 3 N2 (g) | High | 3406 |

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(3)

(Total 7 marks)

**41.** (a) Define the term *average bond enthalpy*.

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(2)

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

C4H8(g) + 6O2(g) → 4CO2(g) + 4H2O(g)

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(3)

(Total 5 marks)

**42.** Given the following data:

C(s) + 2 F2(g) → CF4(g); ∆*H*1= –680 kJ mol–1  
F2(g) → 2F(g); ∆*H*2= +158 kJ mol–1  
C(s) → C(g); ∆*H*3= +715 kJ mol–1

calculate the average bond enthalpy (in kJ mol–1) for the C––F bond.

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(Total 4 marks)

**43.** Two reactions occurring in the manufacture of sulfuric acid are shown below:

reaction I S(s) +O2(g) → SO2(g) Δ*H*Ө = –297 kJ

reaction II SO2(g) + O2(g)  SO3(g) Δ*H*Ө = –92 kJ

(i) State the name of the term Δ*H*Ө. State, with a reason, whether reaction I would be accompanied by a decrease or increase in temperature.

(3)

(ii) At room temperature sulfur trioxide, SO3, is a solid. Deduce, with a reason, whether the Δ*H*Ө value would be more negative or less negative if SO3(s) instead of SO3(g) were formed in reaction II.

(2)

(iii) Deduce the Δ*H*Ө value of this reaction:

S (s) +  O2 (g) → SO3 (g)

(1)

(Total 6 marks)

**44.** (i) Define the term *average bond enthalpy*.

(3)

(ii) Explain why Br2 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:

CH4 (g) + Br2 (g) → CH3Br (g) + HBr (g)

(3)

(iv) Sketch an enthalpy level diagram for the reaction in part (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 (iii):

CH3Br (g) + Br2 (g) → CH2Br2 (g) + HBr (g)

(2)

(Total 11 marks)

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

C4H8 + 6 O2 → 4 CO2 + 4 H2O

(a) Use the data below to calculate the value of Δ*H*Ө for the combustion of but-1-ene.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **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 |

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(3)

(b) State and explain whether the reaction above is endothermic or exothermic.

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(1)

(Total 4 marks)

**46.** Calculate the enthalpy change, Δ*H*4 for the reaction

C + 2H2 + O2 → CH3OH Δ*H*4

using Hess’s Law and the following information.

CH3OH + O2 → CO2 + 2H2O Δ*H*1 = -676 kJ mol-1

C + O2 → CO2 Δ*H*2 = -394 kJ mol-1

H2 + O2 → H2O Δ*H*3 = -242 kJ mol-1

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(Total 4 marks)

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

CH3OH(g) + NH3(g) → CH3NH2(g) + H2O(g)

(a) Define the term *average bond enthalpy*.

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(2)

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

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(4)

(Total 6 marks)

**48.** (a) Define the term *average bond enthalpy*.

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(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

C4H8(g) → 4CO2(g) + 4H2O(g)

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(3)

(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.

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(1)

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



State and explain the relative stabilities of the reactants and products.

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(2)

(Total 8 marks)

**49.** The reaction between ethene and hydrogen gas is exothermic.

(i) Write an equation for this reaction.

(1)

(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)

(Total 5 marks)

**50.** (i) Define the term *average bond enthalpy*.

(2)

(ii) The equation for the reaction of ethyne and hydrogen is:

C2H2(g) + 2H2(g) → C2H6(g)

Use information from Table 10 of the Data Booklet to calculate the change in enthalpy for the reaction.

(2)

(iii) State and explain the trend in the bond enthalpies of the C–Cl, C–Br and C–I bonds.

(2)

(Total 6 marks)

**IB Chemistry SL**

**Topic 5 Answers**

**1.** D

**2.** C

**3.** B

**4.** D

**5.** B

**6.** A

**7.** B

**8.** A

**9.** C

**10.** C

**11.** D

**12.** B

**13.** D

**14.** A

**15.** D

**16.** D

**17.** C

**18.** B

**19.** C

**20.** C

**21.** A

**22.** D

**23.** A

**24.** C

**25.** C

**26.** C

**27.** A

**28.** C

**29.** A

**30.** B

**31.** B

**32.** A

**33.** D

**34.** C

**35.** C

**36.** B

**37.** (a) activation energy = **minimum** energy required for a reaction to occur; 1

(b) curve moved to the right;  
peak lower, 2

Deduct **[1]** if shaded area smaller at T2 or if T2 line touches the x-axis

(c) rate increased;  
as more molecules with energy > *E*a; 2

[5]

**38.** (a)energy for the conversion of a gaseous molecule into (gaseous) atoms;  
(average values) obtained from a number of similar bonds/compounds/*OWTTE*;  
CH4(g) → C(g) + 4H(g); 3

State symbols needed.

(b) (bond breaking) = 1890/654;   
(bond formation) = 2005/769;   
enthalpy = –115(kJ mol–1) 3

Award **[3]** for correct final answer.

Penalize **[1]** for correct answer with wrong sign.

(c) molecules have insufficient energy to react (at room temperature)/  
wrong collision geometry/unsuccessful collisions;   
extra energy needed to overcome the activation energy/*E*a for the reaction; 2

(d)



exothermic shown;

activation energy/Ea shown; 2

[10]

**39.** (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*∆*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;  
∆*H* = (1.463÷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

[10]

**40.** (a) Δ*T* = 23.70 – 23.03 = 0.67 (°C/K); 1

(b)  = 1.281×10–3; 1

(c) (i) Δ*H*c = (*C* Δ*T*)/*n* =  = –5.3×103 kJ mol–1; 1

Use ECF for values of ΔT and n.

(ii) Percentage experimental error == 5.4%; 1

Use ECF for values of ΔHc.

(d) enthalpy change of combustion of sucrose > TNT, and therefore not important;  
rate of reaction for TNT is greater than that of sucrose, so this is valid;  
amount of gas generated (in mol) for sucrose > than that of TNT  
(according to the given equation), so this is not important; 3

[7]

**41.** (a) The amount of energy needed to break 1 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) + (2×348) + (8×412) + (6×496)/7580 (kJ mol–1);  
Bonds made  
(8×743) + (8×463) / 9648 (kJ mol–1);  
Δ*H* = –2068 (kJ mol–1); 3

Award [3] for the correct answer.  
Allow full ECF.  
Allow kJ but no other incorrect units.  
Even if the first two marks are lost, the candidate can score **[1]** for a clear correct subtraction for ΔH.

[5]

**42.** C (s) + 2 F2 (g) → CF4 (g) ∆*H*1 = –680 kJ;  
4 F (g) → 2 F2 (g) ∆*H*2 = 2(–158) kJ;  
C (g) → C (s) ∆*H*3 = –715 kJ;

Accept reverse equations with +∆H values.

C(g) + 4F(g) → CF4(g) ∆*H* = –1711 kJ,

so average bond enthalpy = 

= –428 kJ mol–1; 4

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 2F2 (g) → 4F(g) 2×158 seen.  
Award **[1]** for dividing 1711 or other value by 4.

[4]

**43.** (a) (i) standard enthalpy (change) of reaction;

(temperature) increase;

reaction is exothermic/sign of Δ*H*r 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

[6]

**44.** (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;

(Δ*Hθ* =) –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 | CH4 + Br2  CH3Br + 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

[11]

**45.** (a) (Amount of energy required to break bonds of reactants)

8×412 + 2×348 + 612 + 6×496/7580 (kJ mol-1);

(Amount of energy released during bond formation)

4×2×743 + 4×2×463/9648 (kJ mol-1);

Δ*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** Δ*HӨ* is negative/energy is released; 1

Apply ECF to sign of answer in part (a).

Do not mark if no answer to (a).

[4]

**46.** -1×ΔH1/676;

1×ΔH2/–394;

2×ΔH3/– 484;

Δ*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).

[4]

**47.** (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;

Δ*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]**.

[6]

**48.** (a) amount of energy needed to break one mole of (covalent) bonds;

in the gaseous state;

average calculated from a range of compounds; 2

Award **[1]** each for any two points above.

(b) bonds broken: 161 + 2×348 + 8×412 + 6×496/7580 kJ mol-1;

bonds made: 8×743 + 8×463/9648 kJ mol-1;

(bonds broken - bonds made =) Δ*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

(d) 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

[8]

**49.** (a) (i) C2H4(g) + H2(g) → C2H6(g); 1

State symbols not required for mark

(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

[5]

**50.** (i) energy required to break (a mole of) bonds in the gaseous state  
/energy given out when (a mole of) bonds are made in the  
gaseous state;  
average value from a number of similar compounds; 2

(ii) (*ΔHӨ*reaction = (∑*BE*break - *BE*make))

= [(837) + 2(436)] - [(348 + 4(412)];

= - 287(kJ/kJ mol-1); 2

Award **[1 max]** for 287 or + 287.

(iii) (BE): C-Cl > C-Br > C-I/C-X bond becomes weaker;

halogen size/radius increases/bonding electrons further away from  
the nucleus/bonds become longer; 2

[6]