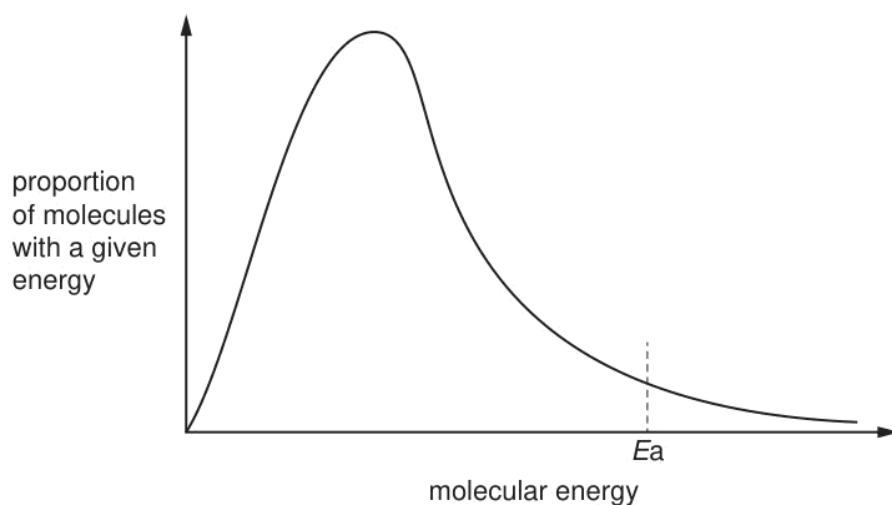


- (b) At a temperature of 700 K a sample of HBr is approximately 10% decomposed. Changing the temperature affects both the rate of decomposition of HBr and the percentage that decomposes.

The Boltzmann distribution for a sample of HBr at 700 K is shown. E_a represents the activation energy for the reaction.



- (i) Using the same axes, sketch a second curve to indicate the Boltzmann distribution at a higher temperature. [2]

(a) The enthalpy change of formation of SO_2 , $\Delta H_f^\circ \text{SO}_2(\text{g})$, is $-296.8 \text{ kJ mol}^{-1}$.

(i) Define the term *enthalpy change of formation*.

.....
.....
..... [2]

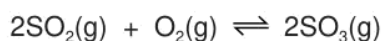
(ii) Use the data to calculate the enthalpy change of formation of $\text{SO}_3(\text{g})$.

$\Delta H_f^\circ \text{SO}_3(\text{g}) = \dots\dots\dots \text{ kJ mol}^{-1}$ [2]

(ii) State the benefit of using a catalyst in this reaction. Explain how it achieves this effect.

.....
.....
.....
..... [2]

(c) At a pressure of $1.50 \times 10^5 \text{ Pa}$, 1.00 mol of sulfur dioxide gas, SO_2 , was mixed with 1.00 mol of oxygen gas, O_2 . The final equilibrium mixture formed was found to contain 0.505 mol of O_2 .



(i) Calculate the amount, in mol, of SO_2 and SO_3 in the equilibrium mixture.

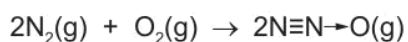
$\text{SO}_2 = \dots\dots\dots \text{ mol}$

$\text{SO}_3 = \dots\dots\dots \text{ mol}$
[1]

(ii) Calculate the partial pressure of oxygen gas, p_{O_2} , in the equilibrium mixture.

$p_{\text{O}_2} = \dots\dots\dots \text{Pa}$ [2]

- (iii) Molecules of N_2O can be formed by the reaction between N_2 and O_2 . The bond between the N and O atoms ($\text{N}\rightarrow\text{O}$) is a co-ordinate (dative covalent) bond.



The enthalpy change of reaction for this reaction is $+82\text{ kJ mol}^{-1}$.

Calculate the bond enthalpy, in kJ mol^{-1} , of the $\text{N}\rightarrow\text{O}$ bond.

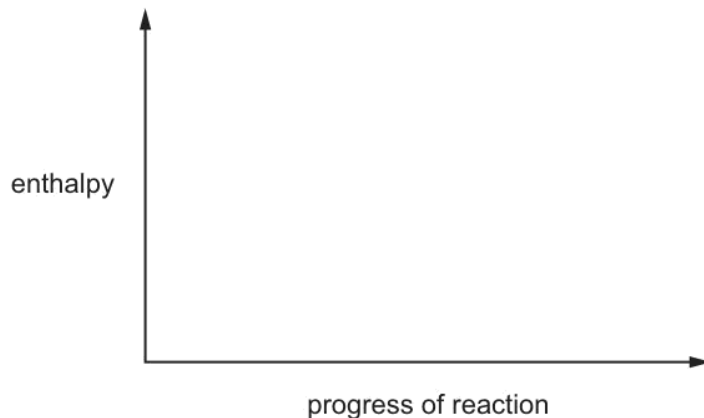
Use relevant data from the *Data Booklet* to answer this question.

bond enthalpy of the $\text{N}\rightarrow\text{O}$ bond = kJ mol^{-1}
[2]

- (iii) HOF is an unstable compound and decomposes to form HF and O₂.



Draw a fully labelled reaction pathway diagram on the axes provided to show the decomposition of HOF into HF and O₂.



[2]

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[Turn over

March 2019 9701/22 · Q2 (c)(iii)

- (iii) During the process of ionisation a sulfur atom loses an electron.



Identify the orbital from which this electron is removed. Explain your answer.

orbital

explanation

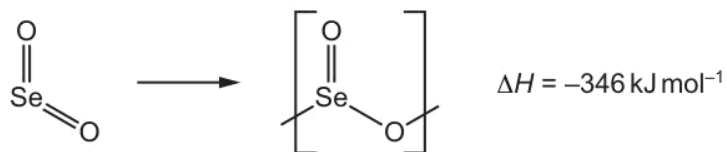
.....

[2]

May/June 2019 9701/23 · Q1 (c)(iii)

- (ii) The most stable oxide of selenium is SeO_2 .

Gaseous SeO_2 reacts to form a solid polymer, as shown. In the reaction one $\text{Se}=\text{O}$ is replaced by two $\text{Se}-\text{O}$ to form a polymer.



The bond enthalpy of $\text{Se}=\text{O}$ is 514 kJ mol^{-1} .

Use these data to calculate the bond enthalpy, in kJ mol^{-1} , of $\text{Se}-\text{O}$.

bond enthalpy of $\text{Se}-\text{O} = \dots\dots\dots \text{kJ mol}^{-1}$
[2]

- (ii) A student carries out an experiment to determine the enthalpy change of combustion of $\text{C}_4\text{H}_4\text{S(l)}$.

Explain the meaning of the term *enthalpy change of combustion*.

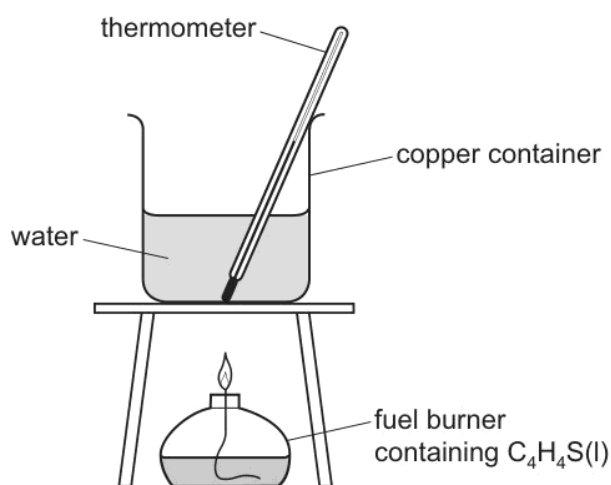
.....
.....
..... [2]

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- (iii) The student uses the following apparatus in the experiment.



mass of water in copper container/g	200
initial temperature of water/ $^{\circ}\text{C}$	18.5
highest temperature of water/ $^{\circ}\text{C}$	37.5

Calculate the heat energy released, in J, by the reaction.

Assume that 4.18 J of heat energy changes the temperature of 1.0 cm^3 of water by 1.0°C .

Assume no heat is lost to the surroundings.

heat energy released = J
[2]

- (iv) The student used 0.63 g of $\text{C}_4\text{H}_4\text{S(l)}$ in the experiment.

Calculate the enthalpy change of combustion of thiophene, $\Delta H_c(\text{C}_4\text{H}_4\text{S(l)})$. Include a sign in your answer.

$$\Delta H_c(\text{C}_4\text{H}_4\text{S(l)}) = \dots\dots\dots \text{kJ mol}^{-1}$$

[2]

[Total: 13]

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[Turn over

8

Oct/Nov 2019 9701/21 · Q3 (b)(ii), (b)(iii), (b)(iv) (cont.)

(ii) Define the term *enthalpy change of formation*.

.....

.....

..... [2]

(iii) Use the data given to calculate the enthalpy change of formation of $\text{MgO}_2(\text{s})$.

$$\Delta H_f^\circ \text{MgO}_2(\text{s}) = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

- (iv) Magnesium peroxide decomposes slowly to form magnesium oxide and oxygen.



Use your answer to (g)(iii) and the data in the table to calculate the enthalpy change of this reaction.

If you were unable to obtain an answer to (g)(iii), use the value $\Delta H_f = -550 \text{ kJ mol}^{-1}$. This is **not** the correct answer.

enthalpy change of reaction = kJ mol^{-1} [1]

[Total: 19]

March 2020 9701/22 · Q1 (g)(ii), (g)(iii), (g)(iv) (cont.)

- (a) Explain what is meant by the term *enthalpy change of formation*.

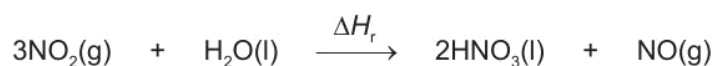
.....

 [2]

- (b) Complete the Hess' cycle using the values given in the table and hence calculate the enthalpy change, ΔH_r , for this reaction.

Show your working.

substance	$\Delta H_f / \text{kJ mol}^{-1}$
$\text{NO}_2(\text{g})$	34.0
$\text{H}_2\text{O}(\text{l})$	−286
$\text{HNO}_3(\text{l})$	−173
$\text{NO}(\text{g})$	91.1



$$\Delta H_r = \dots\dots\dots \text{kJ mol}^{-1}$$

[3]

- (c) Nitrogen and oxygen do not react at normal atmospheric temperatures.

Explain why.

.....

.....

.....

..... [2]

Nitrogen oxides can be formed naturally in the Earth's atmosphere from nitrogen and oxygen in the air.

- (d) State **one** way that nitrogen oxides are produced naturally.

..... [1]

- (e) Nitrogen dioxide, NO_2 , acts as a homogeneous catalyst in the oxidation of atmospheric sulfur dioxide.

- (i) Explain why NO_2 is described as a homogeneous catalyst.

.....

.....

.....

..... [3]

- (ii) Write equations which describe the two reactions occurring when NO_2 acts as a catalyst in the formation of sulfur trioxide from sulfur dioxide.

.....

..... [2]

[Total: 13]

May/June 2020 9701/21 · Q3 (a), (b), (c), (d), (e)(i), (e)(ii) (cont.)

- (d) 1.00 g of sucrose, $C_{12}H_{22}O_{11}$, is completely combusted. The heat energy produced is used to increase the temperature of 250 g of water inside a calorimeter from 25.0 °C to 40.7 °C.

These data can be used to calculate the enthalpy change of combustion of sucrose.

- (i) Explain what is meant by the term *enthalpy change of combustion of sucrose*.

.....
.....
..... [2]

- (ii) Use the *Data Booklet* to calculate the enthalpy change, in kJ mol^{-1} , for the combustion of sucrose.

Assume that all of the heat energy produced is transferred to the water.

Show your working.

enthalpy change of combustion of sucrose = kJ mol^{-1}
[3]

[Total: 12]

May/June 2020 9701/22 · Q3 (d)(i), (d)(ii)

- (c) (i) Explain what is meant by the term *enthalpy change of combustion*.

.....
..... [2]

- (ii) Write the equation for the complete combustion of sucrose.

..... [1]

The enthalpy change of reaction **Z**, ΔH_r , can be calculated using the enthalpy change of combustion data given in the table.

substance	enthalpy change of combustion, $\Delta H_c / \text{kJ mol}^{-1}$
sucrose	−5643
glucose	−2805
fructose	−2810

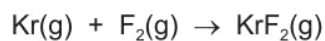
May/June 2020 9701/23 · Q3 (c)(i), (c)(ii)

(ii) Use the data to calculate the enthalpy change of reaction 2.

compound	$\Delta H_f / \text{kJ mol}^{-1}$
CaO(s)	−635
SO ₂ (g)	−297
CaSO ₄ (s)	−1434

enthalpy change of reaction 2 = kJ mol^{−1} [2]

- (c) Krypton reacts with fluorine in the presence of ultraviolet light to make krypton difluoride, $\text{KrF}_2(\text{g})$.

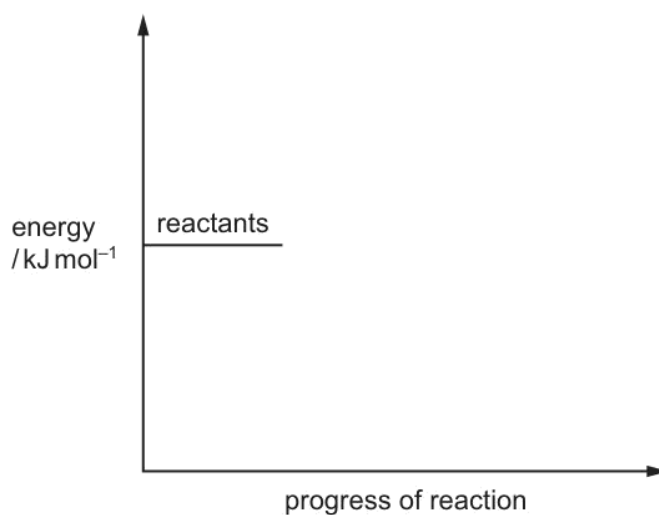


activation energy for the reaction, $E_a = +385 \text{ kJ mol}^{-1}$

enthalpy change of formation of KrF_2 , $\Delta H_f = +60.2 \text{ kJ mol}^{-1}$

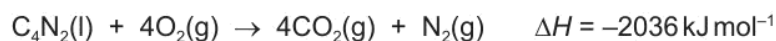
- (i) Use this information to complete the reaction profile diagram for the formation of KrF_2 . Label E_a and ΔH_f on the diagram.

Assume the reaction proceeds in one step.



[2]

- (a) The equation for the complete combustion of **P**, $\text{C}_4\text{N}_2(\text{l})$, is shown.



- (i) The enthalpy change of formation, ΔH_f , of $\text{CO}_2(\text{g})$ is -384 kJ mol^{-1} .

Calculate the enthalpy change of formation, ΔH_f , of **P**, in kJ mol^{-1} .

ΔH_f of **P** = kJ mol^{-1} [2]

- (ii) One of the products of the complete combustion of **P** is nitrogen gas, $\text{N}_2(\text{g})$.

Explain the lack of reactivity of nitrogen.

..... [1]

(b) **Q** forms when HCN reacts with ethyne, $\text{H}-\text{C}\equiv\text{C}-\text{H}$.

(i) Ethyne, HCN and **Q** are all weak Brønsted–Lowry acids.

Explain what is meant by the term *weak Brønsted–Lowry acid*.

.....
.....
..... [2]

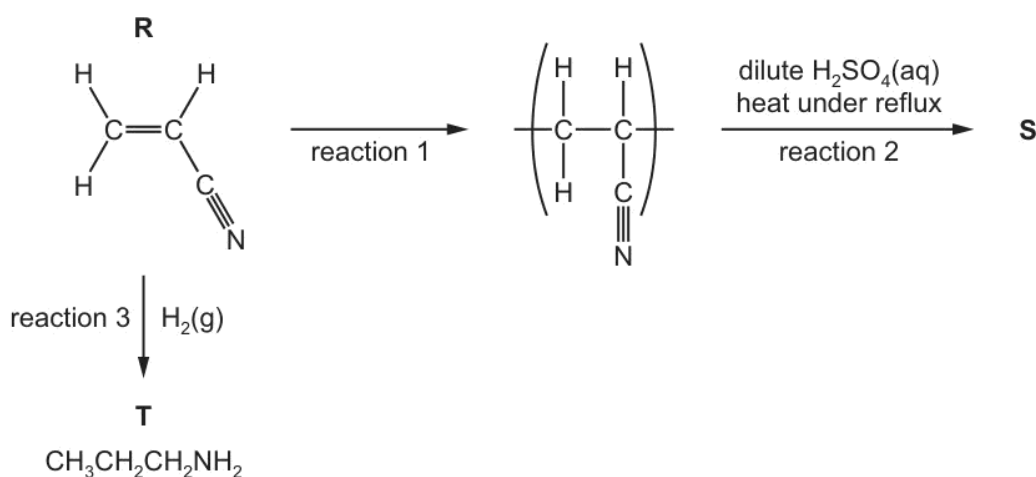
(ii) Ethyne, HCN and **Q** all contain triple bonds between two atoms.

A triple bond consists of one sigma (σ) and two pi (π) bonds.

Draw a labelled diagram to show the formation of one pi (π) bond.

[2]

(d) The flow chart shows some reactions of **R**.



(i) Name the type of reaction shown in reaction 1.

..... [1]

(ii) Draw the structure of **S**, the organic product of reaction 2.

[1]

(iii) Name **T**.

..... [1]

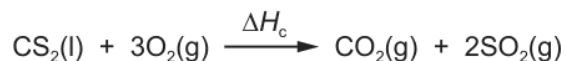
(iv) **T** can also be formed by the reaction of $\text{CH}_3\text{CH}_2\text{CH}_2\text{Br}$ with ammonia.

State the necessary conditions of this reaction.

..... [1]

[Total: 13]

(b) The enthalpy change of combustion of $\text{CS}_2(\text{l})$ is represented by the following equation.



(i) Define *enthalpy change of combustion*.

.....
.....
..... [2]

(ii) The table shows the enthalpy changes of formation of $\text{CS}_2(\text{l})$, $\text{CO}_2(\text{g})$ and $\text{SO}_2(\text{g})$.

compound	enthalpy change of formation, $\Delta H_f^\circ / \text{kJ mol}^{-1}$
$\text{CS}_2(\text{l})$	+89.7
$\text{CO}_2(\text{g})$	−394
$\text{SO}_2(\text{g})$	−297

Use the data in the table to calculate the enthalpy change of combustion, ΔH_c , of $\text{CS}_2(\text{l})$, in kJ mol^{-1} .

Show your working.

ΔH_c of $\text{CS}_2(\text{l})$ = kJ mol^{-1}
[2]

Oct/Nov 2021 9701/21 · Q1 (b)(i), (b)(ii)

- (b) The standard enthalpy change of formation, $\Delta H_f^\ominus$, of HI(g) is $+26.5 \text{ kJ mol}^{-1}$.

Define the term *standard enthalpy change of formation*.

.....

.....

..... [2]

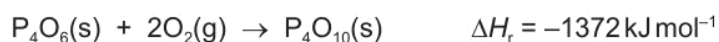
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3

Oct/Nov 2021 9701/22 · Q1 (b)

- (iii) When $\text{P}_4\text{O}_6(\text{s})$ is heated with oxygen it forms $\text{P}_4\text{O}_{10}(\text{s})$.



The enthalpy change of formation, ΔH_f , of $\text{P}_4\text{O}_{10}(\text{s})$ is $-3012 \text{ kJ mol}^{-1}$.

Calculate the enthalpy change of formation, ΔH_f , of $\text{P}_4\text{O}_6(\text{s})$.

ΔH_f of $\text{P}_4\text{O}_6(\text{s}) = \dots\dots\dots \text{ kJ mol}^{-1}$ [1]

March 2022 9701/22 · Q2 (c)(iii)

- (iii) Define enthalpy change of formation.

.....
..... [2]

- (iv) When 3.645 g of Mg(s) burns in excess N₂(g) to form Mg₃N₂(s), 23.05 kJ of energy is released.

Calculate the enthalpy change of formation, ΔH_f , of Mg₃N₂. Show your working.

$$\Delta H_f (\text{Mg}_3\text{N}_2) = \dots\dots\dots [3]$$

[Total: 9]

3

Question 2 starts on the next page.

4

May/June 2022 9701/22 · Q1 (b)(iii), (b)(iv) (cont.)

- (ii) Calculate the enthalpy change of formation, ΔH_f , of $\text{SCl}_2(\text{l})$. You may find it useful to use Hess's Law to construct an energy cycle.

enthalpy change of formation of $\text{SCl}_2(\text{l})$, $\Delta H_f = \dots\dots\dots \text{kJ mol}^{-1}$
[2]

Oct/Nov 2022 9701/23 · Q3 (d)(ii)

- (a) Define enthalpy change of neutralisation, ΔH_{neut} .

.....

 [2]

- (b) An experiment is carried out to calculate ΔH_{neut} for the reaction between NaOH(aq) and H₂SO₄(aq).

100 cm³ of 1.00 mol dm⁻³ NaOH(aq) is added to 75 cm³ of 1.00 mol dm⁻³ H₂SO₄(aq) in a polystyrene cup and stirred. Results from the experiment are shown in Table 2.1.

Table 2.1

initial temperature of NaOH(aq) / °C	20.0
initial temperature of H ₂ SO ₄ (aq) / °C	20.0
maximum temperature of mixture / °C	27.8

- (i) Use equation 1 to calculate the amount, in mol, of H₂SO₄(aq) that is neutralised in the experiment.

amount of H₂SO₄(aq) neutralised = mol [1]

- (ii) Calculate ΔH_{neut} using the results in Table 2.1. Include units in your answer.

Assume that:

- the specific heat capacity of the final solution is 4.18 J g⁻¹ K⁻¹
- 1.00 cm³ of the final solution has a mass of 1.00 g
- there is no heat loss to the surroundings
- full dissociation of H₂SO₄(aq) occurs
- the experiment takes place at constant pressure.

Show your working.

ΔH_{neut} = units [3]

- (c) (i) Complete the equation for the reaction that occurs when a solution of $\text{Ba}(\text{OH})_2$ is added to aqueous sulfuric acid. Include state symbols.

..... $\text{H}_2\text{SO}_4(\text{aq})$ + $\text{Ba}(\text{OH})_2(\text{aq}) \rightarrow$ [2]

- (ii) Suggest why the enthalpy change of neutralisation cannot be determined using the addition of dilute sulfuric acid to aqueous barium hydroxide.

.....
..... [1]

[Total: 9]

May/June 2023 9701/21 · Q2 (a), (b)(i), (b)(ii), (c)(ii) (cont.)

- (a) Define enthalpy change of neutralisation, ΔH_{neut} .

.....
.....
..... [2]

- (b) In an experiment, 50.0 cm³ of 2.00 mol dm⁻³ NaOH(aq) is added to 60.0 cm³ of 1.00 mol dm⁻³ H₂SO₄(aq) in a polystyrene cup and stirred. Both solutions have a temperature of 21.4 °C before mixing. The maximum temperature of the mixture is measured.

- (i) Use equation 1 to calculate the amount, in mol, of H₂SO₄(aq) that is neutralised in the experiment.

amount of H₂SO₄ neutralised = mol [1]

- (ii) Calculate the theoretical maximum temperature of the mixture in this experiment.

Assume that:

- enthalpy change of neutralisation, ΔH_{neut} , of NaOH(aq) and H₂SO₄(aq) is -57.1 kJ mol⁻¹
- full dissociation of H₂SO₄(aq) occurs
- the specific heat capacity of the final solution is 4.18 J g⁻¹ K⁻¹
- 1.00 cm³ of the final solution has a mass of 1.00 g
- there is no heat loss to the surroundings
- the experiment takes place at constant pressure.

Show your working.

theoretical maximum temperature = °C [3]

(c) The enthalpy change of neutralisation of $\text{CH}_3\text{COOH}(\text{aq})$ and $\text{NaOH}(\text{aq})$ is $-55.2 \text{ kJ mol}^{-1}$.

(i) Complete the equation for the reaction.

..... CH_3COOH + NaOH → [1]

(ii) Values for the enthalpy change of neutralisation, ΔH_{neut} , are shown in Table 3.1.

Table 3.1

reagents	$\Delta H_{\text{neut}}/\text{kJ mol}^{-1}$
$\text{NaOH} + \text{HCl}$	-57.1
$\text{NaOH} + \text{CH}_3\text{COOH}$	-55.2

Suggest why the value for ΔH_{neut} of the weak acid, CH_3COOH , reacting with NaOH is different to the value obtained using the strong acid, HCl . Assume that the values are determined under the same conditions.

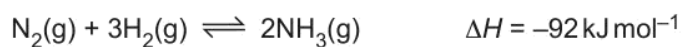
.....
..... [1]

[Total: 8]

May/June 2023 9701/23 · Q3 (a), (b)(i), (b)(ii), (c)(i), (c)(ii) (cont.)

(c) The Haber process involves the reaction of N_2 and H_2 to form ammonia, NH_3 .

A catalyst is used, which allows the process to be carried out at a lower temperature and pressure.



(i) Use the information in (c) to complete Table 3.1.

Table 3.1

compound	enthalpy change of formation, $\Delta H_f / \text{kJ mol}^{-1}$
N_2	
H_2	
NH_3	

[2]

March 2024 9701/22 · Q3 (c)(i)

- 4 (a) Define enthalpy change of formation.

.....

 [2]

- (b) Iron is made when iron(III) oxide is heated with carbon monoxide, as shown by reaction 2.

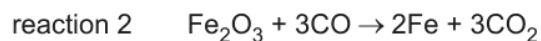


Table 4.1 shows enthalpy change of formation data measured at 298 K and 101 kPa.

Table 4.1

substance	equation	value for $\Delta H_f^\ominus / \text{kJ mol}^{-1}$
Fe_2O_3		–824.2
CO		–110.5
CO_2	$\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$	–393.5

- (i) Complete Table 4.1 by adding equations with relevant state symbols to represent:

- standard enthalpy change of formation for Fe_2O_3
- standard enthalpy change of formation for CO.

[2]

- (ii) Use the data in Table 4.1 to calculate the enthalpy change of reaction, ΔH_r , in kJ mol^{-1} , for reaction 2.

Show your working.

$$\Delta H_r = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

[Total: 6]

May/June 2024 9701/21 · Q4 (a), (b)(i), (b)(ii) (cont.)

- 3 (a) Write an equation to show the reaction for the standard enthalpy change of formation of H_2O . Include state symbols.

..... [2]

- (b) Water is one of the products in the reaction of B_2O_3 and NH_3 , as shown in reaction 2.

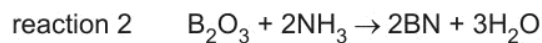


Table 3.1 shows information about the standard enthalpy change of formation, $\Delta H_f^\ominus$, of some substances.

Table 3.1

substance	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
B_2O_3	–1264
NH_3	–46
BN	–134
H_2O	–286

Calculate the enthalpy change, ΔH , for reaction 2 using the data from Table 3.1.

$\Delta H = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

(c) Boron carbide is a hard crystalline solid that has a melting point greater than 2000 °C.

(i) Suggest the structure and bonding in boron carbide.

..... [1]

(ii) 100 g of pure boron carbide contains 78.26 g of boron.

Calculate the empirical formula of boron carbide.

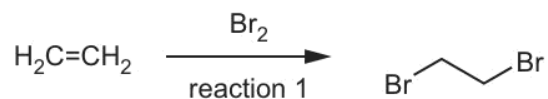
Show your working.

empirical formula of boron carbide [2]

[Total: 7]

May/June 2024 9701/23 · Q3 (a), (b), (c)(i), (c)(ii) (cont.)

(b) The enthalpy change of reaction 1, $\Delta H_r = -90.0 \text{ kJ mol}^{-1}$.



The enthalpy change of formation of ethene, $\Delta H_f = +52.2 \text{ kJ mol}^{-1}$.

Calculate the enthalpy change of formation of 1,2-dibromoethane.

ΔH_f of 1,2-dibromoethane = kJ mol^{-1} [1]

Oct/Nov 2024 9701/22 · Q4 (b)

- (ii) Use the data in Fig. 3.2 and in Table 3.1 to calculate the enthalpy change of reaction 2, ΔH_2 .

Table 3.1

compound	enthalpy change of combustion, $\Delta H_c / \text{kJ mol}^{-1}$
	−3292
	−3115
H_2	−286

$$\Delta H_2 = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

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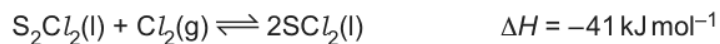
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10

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- (iii) $\text{S}_2\text{Cl}_2(\text{l})$ is yellow and $\text{SCl}_2(\text{l})$ is red.



State what is observed when the following changes are made to an equilibrium mixture of $\text{S}_2\text{Cl}_2(\text{l})$ and $\text{SCl}_2(\text{l})$.

Explain your answers.

- The equilibrium mixture is warmed gently.

observation

explanation

.....

.....

- The overall pressure of the equilibrium mixture is increased.

observation

explanation

.....

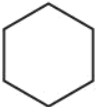
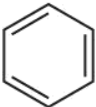
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[4]

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- (ii) Use the data in Fig. 3.2 and in Table 3.1 to calculate the enthalpy change of reaction 2, ΔH_2 .

Table 3.1

compound	enthalpy change of formation, $\Delta H_f^\circ/\text{kJ mol}^{-1}$
	-156
	+48

$\Delta H_2 = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

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- (ii) Use the data in Fig. 3.2 and in Table 3.1 to calculate the enthalpy change of reaction 2, ΔH_2 .

Table 3.1

compound	enthalpy change of combustion, $\Delta H_c / \text{kJ mol}^{-1}$
	−3292
	−3115
H_2	−286

$$\Delta H_2 = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

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[Turn over]

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(iii) Define enthalpy change of formation, ΔH_f° .

.....

.....

..... [2]

(iv) Table 1.1 gives data relevant to the reaction of $\text{Cl}_2(\text{g})$ with $\text{NH}_3(\text{g})$.

Table 1.1

compound	enthalpy change of formation, $\Delta H_f^\circ / \text{kJ mol}^{-1}$
$\text{NH}_3(\text{g})$	−46
$\text{HCl}(\text{g})$	−92

Use the data in Table 1.1 to calculate the enthalpy change of the reaction of $\text{Cl}_2(\text{g})$ with $\text{NH}_3(\text{g})$.

enthalpy change of reaction = kJ mol^{-1} [2]