

Please write clearly in block capitals.

Centre number

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Candidate number

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Surname *StemUp Model Solution*

Forename(s)

Candidate signature

INTERNATIONAL A-LEVEL CHEMISTRY (9620)

Unit 3: Inorganic 2 and Physical 2

Friday 18 January 2019 07:00 GMT Time allowed: 1 hour 30 minutes

Materials

For this paper you must have:

- the Periodic Table/Data Sheet, provided as an Insert
- a ruler with millimetre measurements
- a scientific calculator, which you are expected to use where appropriate.

Instructions

- Use black ink or black ball-point pen.
- Fill in the boxes at the top of this page.
- Answer **all** questions.
- You must answer the questions in the spaces provided. Do not write outside the box around each page or on blank pages
- If you need extra space for your answer(s), use the lined pages at the end of this book. Write the question number against your answer(s).
- All working must be shown.
- Do all rough work in this book. Cross through any work you do not want to be marked.

Information

- The marks for questions are shown in brackets.
- The maximum mark for this paper is 80.

For Examiner's Use	
Question	Mark
1	
2	
3	
4	
5	
6	
7	
8	
9	
TOTAL	



Answer **all** questions in the spaces provided.

Do not write
outside the
box

0 1

This question is about compounds of the elements in Period 3.

0 1 . 1

Phosphorus burns in air to form phosphorus(V) oxide.

Write an equation for this reaction.

[1 mark]



0 1 . 2

Sulfur forms two oxides, SO_2 and SO_3

Both oxides react with water to form acidic solutions.

State the formula of each acid formed and the approximate pH of a 0.1 mol dm^{-3} solution of each acid.

[2 marks]

Formula of acid formed by SO_2 H_2SO_3

pH 3

Formula of acid formed by SO_3 H_2SO_4

pH 1

0 1 . 3

Aluminium oxide can act as an acid and can act as a base.

Write an equation to show aluminium oxide acting as an acid.

[1 mark]



0 1 . 4

Silicon reacts with oxygen to form SiO_2 and with chlorine to form SiCl_4

Table 1 shows the melting points of SiO_2 and SiCl_4

Table 1

Compound	Melting point / °C
SiO_2	1710
SiCl_4	-69

Explain, in terms of structure and bonding, why SiO_2 has a very high melting point but SiCl_4 has a very low melting point.

[4 marks]

SiO_2 *SiO_2 has a giant covalent (macromolecular) lattice structure.*

To melt it, many strong covalent bonds between the Si and O atoms must be broken, which requires a lot of energy.

SiCl_4 *SiCl_4 has a simple molecular structure.*

Only weak van der Waals' (intermolecular) forces between the molecules must be overcome, which requires little energy.

8

Turn over for the next question

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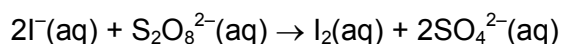


0 2

This question is about Fe^{2+} and Fe^{3+} ions.

0 2 . 1

The reaction between I^- ions and $\text{S}_2\text{O}_8^{2-}$ ions has a high activation energy and is slow unless a catalyst is used.



Fe^{2+} ions catalyse the reaction between I^- ions and $\text{S}_2\text{O}_8^{2-}$ ions.

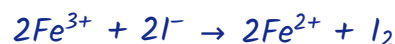
Write **two** equations to show how Fe^{2+} ions catalyse this reaction.

[2 marks]

Equation 1



Equation 2



0 2 . 2

Describe how Fe(II) enables oxygen to be transported in the blood.

[2 marks]

The iron(II) is in haemoglobin in red blood cells.

*An oxygen molecule forms a co-ordinate (dative) bond to the Fe^{2+} ion,
so O_2 is carried and can be released.*

0 2 . 3

State why carbon monoxide is toxic.

[1 mark]

*CO bonds more strongly to the Fe^{2+} in haemoglobin than O_2 does, so it
takes the place of oxygen and oxygen cannot be transported.*



0 2 . 4

Iron also forms the Fe^{3+} ion.

A student adds sodium carbonate solution to a solution containing $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$. Effervescence is observed.

State **one** other observation the student would make.

Write an equation for the reaction between $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$ and carbonate ions.

[2 marks]

Observation a brown precipitate forms

Equation



0 2 . 5

Explain why $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$ ions are more acidic than $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ ions.

[2 marks]

Fe^{3+} has a higher charge (a higher charge to size ratio) than Fe^{2+} , so it is more polarising.

It pulls electron density out of the O-H bonds in the water ligands, weakening them, so H^+ ions are released more readily.

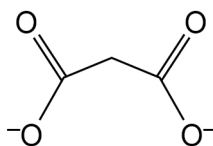
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0 3

A propanedioate ion can act as a bidentate ligand.
The skeletal formula of a propanedioate ion is shown.



0 3 . 1

Give the meaning of the term bidentate ligand.

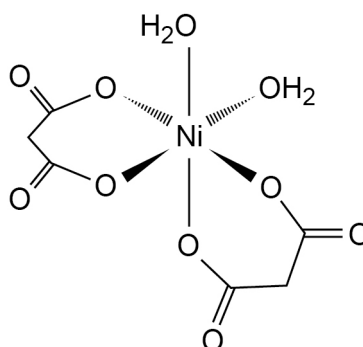
[1 mark]

A ligand that donates two electron pairs to a metal ion, from two different atoms, forming two co-ordinate bonds.

0 3 . 2

Ni^{2+} ions form a complex ion **Z** that contains two propanedioate ligands and two water ligands.

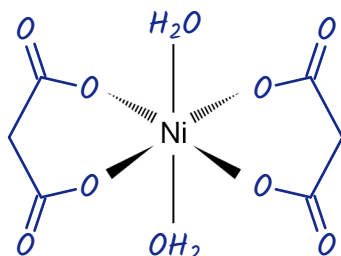
Z exists as three stereoisomers. One of these stereoisomers is shown. The charge has not been shown.



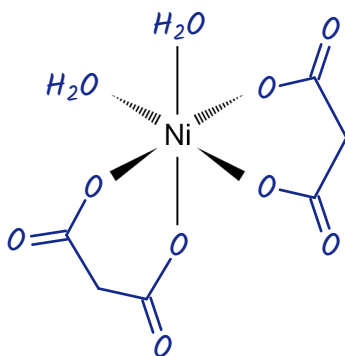
Complete the diagrams of the other two stereoisomers of **Z**.
You do **not** need to include any charges on your diagrams.

For each stereoisomer, place one or more ticks in the boxes to identify the type or types of isomer shown.

[4 marks]



<i>cis</i>	
<i>trans</i>	✓
optical	



<i>cis</i>	✓
<i>trans</i>	
optical	✓

5

Turn over for the next question

Turn over ►



0 4

A student investigates the solubility of hydrated iron(II) ethanedioate ($\text{FeC}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$) in water.

The student uses the following method.

Step 1 Add hydrated iron(II) ethanedioate to approximately 500 cm^3 of water at 25°C

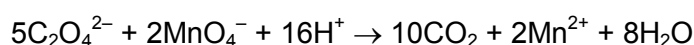
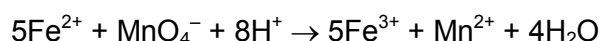
Step 2 Stir until no more solid dissolves and a saturated solution is formed.

Step 3 Remove a 25.0 cm^3 portion of the saturated solution. Add 25.0 cm^3 of dilute sulfuric acid and heat to about 60°C

Step 4 Titrate the hot solution with acidified potassium manganate(VII) solution.

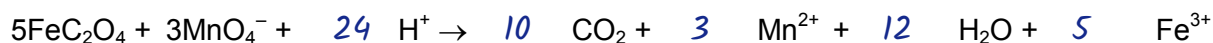
0 4 . 1

The Fe^{2+} and $\text{C}_2\text{O}_4^{2-}$ ions react with MnO_4^- ions in acidic solution as shown.



Complete the overall equation for the reaction between FeC_2O_4 and MnO_4^- ions in acidic solution.

[1 mark]



0 4 . 2

Suggest why the solution is heated in **Step 3**.

[1 mark]

The reaction is slow at room temperature (it has a high activation energy), so heating increases the rate of reaction.

0 4 . 3

During **Step 3** some evaporation of the solution occurs.

State and explain the effect, if any, this has on the volume of potassium manganate(VII) solution needed in **Step 4**.

[2 marks]

Effect on volume of potassium manganate(VII) *no effect*

Explanation *Only water is lost by evaporation, so the amount (moles) of iron(II) ethanedioate in the flask does not change.*

The same amount of MnO_4^- is still needed to react with it, so the titre is unchanged.



0 4 . 4

Explain why no indicator is needed for this titration.
State the colour at the end point.

[2 marks]

Explanation Once all the iron(II) ethanedioate has reacted, excess purple

MnO_4^- remains and colours the solution; self indicating.

Colour at end point (pale) pink

0 4 . 5

In **Step 4**, the concentration of potassium manganate(VII) is $6.40 \times 10^{-3} \text{ mol dm}^{-3}$
The mean titre is 12.45 cm^3

Calculate the solubility, in g dm^{-3} , of hydrated iron(II) ethanedioate at 25°C

[5 marks]

$$n(\text{MnO}_4^-) = 6.40 \times 10^{-3} \times 12.45 \div 1000 = 7.968 \times 10^{-5} \text{ mol}$$

$$n(\text{FeC}_2\text{O}_4) = 7.968 \times 10^{-5} \times 5/3 = 1.328 \times 10^{-4} \text{ mol}$$

$$n(\text{FeC}_2\text{O}_4) \text{ in } 1 \text{ dm}^3 = 1.328 \times 10^{-4} \times 1000/25.0 = 5.312 \times 10^{-3} \text{ mol}$$

$$M_r(\text{FeC}_2\text{O}_4 \cdot 2\text{H}_2\text{O}) = 55.8 + 24.0 + 64.0 + 36.0 = 179.8$$

$$\text{solubility} = 5.312 \times 10^{-3} \times 179.8 = 0.955 \text{ g dm}^{-3}$$

Solubility 0.955 g dm^{-3}

Question 4 continues on the next page

Turn over ►



0 4 . 6

Iron(II) ethanedioate solution is yellow.

Outline a method, using a colorimeter, to determine the concentration of a solution of iron(II) ethanedioate.

[2 marks]

Make up solutions of iron(II) ethanedioate of known concentrations.

Using a suitable (blue) filter and zeroing with water, measure the absorbance of each.

Plot a calibration graph of absorbance against concentration.

Measure the absorbance of the unknown solution and read its concentration from the calibration graph.

13



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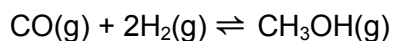
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0 5

Methanol can be made by reacting carbon monoxide with hydrogen.



0 5 . 1

Write an expression for K_p for the manufacture of methanol from carbon monoxide and hydrogen.

[1 mark]

$$K_p = \frac{p(\text{CH}_3\text{OH})}{p(\text{CO}) \times p(\text{H}_2)^2}$$

0 5 . 2

What are the units of K_p for this equilibrium?Tick (✓) **one** box.

[1 mark]

kPa²☐kPa⁻²☒mol² dm⁻⁶☐mol⁻² dm⁶☐

0 5 . 3

Explain why the value of K_p for this reaction decreases when the temperature is increased.

[2 marks]

The forward reaction is exothermic.

Increasing the temperature moves the position of equilibrium to the left

(the endothermic direction) to oppose the increase in temperature, so

the partial pressure of CH₃OH falls and K_p decreases.



A mixture of 10.0 mol of each of CO, H₂ and CH₃OH is placed in a flask of volume 2.50 dm³ and left to reach equilibrium at a given temperature.

At equilibrium, the flask contains 6.0 mol of CO

The total pressure in the flask is 5000 kPa

0 5 . 4

Calculate the amount, in mol, of H₂ and of CH₃OH in the flask at equilibrium.

[2 marks]

$$CO \text{ reacted} = 10.0 - 6.0 = 4.0 \text{ mol}$$

$$H_2 \text{ reacted} = 2 \times 4.0 = 8.0 \text{ mol, so } H_2 \text{ left} = 10.0 - 8.0 = 2.0 \text{ mol}$$

$$CH_3OH \text{ formed} = 4.0 \text{ mol, so } CH_3OH = 10.0 + 4.0 = 14.0 \text{ mol}$$

Amount of H₂ 2.0 mol

Amount of CH₃OH 14.0 mol

0 5 . 5

Calculate the value of K_p under these conditions. You do **not** need to give units for your answer in this part of the question.

(If you were unable to answer question **05.4** you should use 8.0 mol as the amount of H₂ and 8.0 mol as the amount of CH₃OH. These are **not** the correct values.)

[3 marks]

$$\text{total moles} = 6.0 + 2.0 + 14.0 = 22.0$$

$$\text{mole fractions: } CO = 0.273, H_2 = 0.0909, CH_3OH = 0.636$$

$$p(CO) = 0.2727 \times 5000 = 1364 \text{ kPa, } p(H_2) = 0.0909 \times 5000 = 454.5 \text{ kPa}$$

$$p(CH_3OH) = 0.6364 \times 5000 = 3182 \text{ kPa}$$

$$K_p = 3182 \div (1364 \times 454.5^2) = 1.13 \times 10^{-5}$$

K_p 1.13×10^{-5}

9

Turn over for the next question

Turn over ►



0 6

Table 2 shows the standard electrode potentials ($E^\ominus$) for two half-cells.

Table 2

Half-equation	$E^\ominus / \text{V}$
$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Cr}^{2+}(\text{aq})$	-0.41
$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Co}(\text{s})$	-0.28

0 6

1

A standard electrode potential is measured by comparison with a standard hydrogen electrode.

State the temperature and pressure used in a standard hydrogen electrode.

[1 mark]

Temperature 298 K (25 °C)

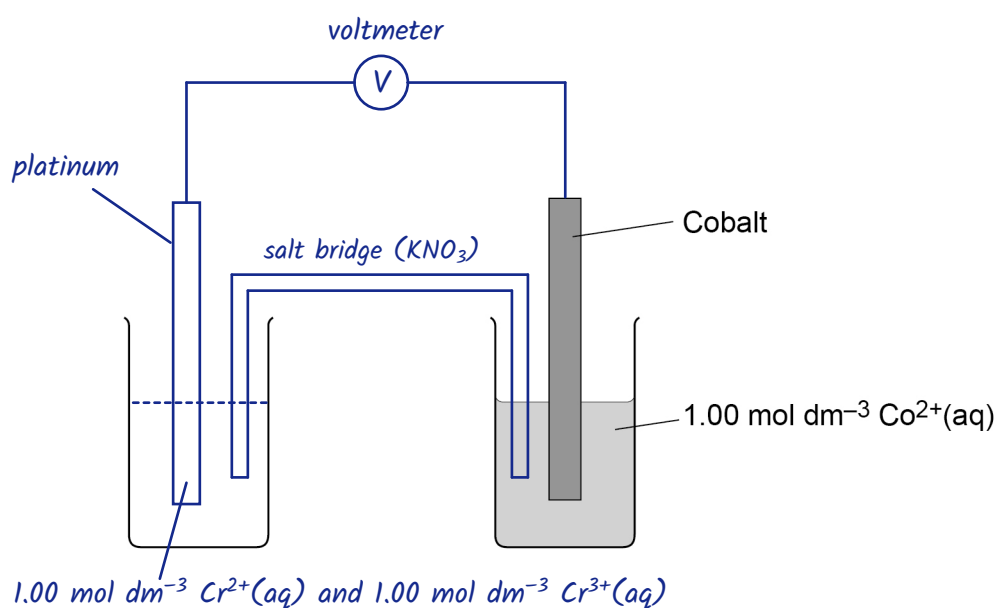
Pressure 100 kPa

An electrochemical cell is set up using the two half-cells from **Table 2**.
The EMF of the cell is measured.

0 6

2

Complete and label the diagram in **Figure 1** to show the electrochemical cell formed from these two half-cells.
Include in your diagram the piece of apparatus needed to measure the EMF of the cell.

[4 marks]**Figure 1**

06.3

Identify the negative electrode in the cell in question **06.2**.
Use the data in **Table 2** to give a reason for your answer.

[1 mark]

Negative electrode platinum (the $\text{Cr}^{3+}/\text{Cr}^{2+}$ half-cell)

Reason The $\text{Cr}^{3+}/\text{Cr}^{2+}$ half-cell has the more negative electrode potential (-0.41 V), so it is the better reducing agent and loses electrons.

06.4

Deduce the EMF of the cell in question **06.2**.

[1 mark]

EMF 0.13 V

06.5

Write the conventional representation of the cell in question **06.2**.

[1 mark]

$\text{Pt} \mid \text{Cr}^{2+}(\text{aq}), \text{Cr}^{3+}(\text{aq}) \parallel \text{Co}^{2+}(\text{aq}) \mid \text{Co}(\text{s})$

Question 6 continues on the next page

Turn over ►



Table 3 shows the standard electrode potential values ($E^\ominus$) for some half-cells.

Table 3

Half-equation	$E^\ominus / \text{V}$
$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg}(\text{s})$	-2.38
$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Cr}^{2+}(\text{aq})$	-0.41
$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Co}(\text{s})$	-0.28
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+1.33
$\text{HOI}(\text{aq}) + \text{H}^+(\text{aq}) + \text{e}^- \rightarrow \frac{1}{2}\text{I}_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+1.45

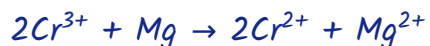
0 6 . 6 $\text{Cr}^{3+}(\text{aq})$ can act as an oxidising agent and as a reducing agent.

Use data from **Table 3** to write:

- **one** equation for a reaction where $\text{Cr}^{3+}(\text{aq})$ acts as an oxidising agent
- **one** equation for a reaction where $\text{Cr}^{3+}(\text{aq})$ acts as a reducing agent.

[2 marks]

$\text{Cr}^{3+}(\text{aq})$ as an oxidising agent



$\text{Cr}^{3+}(\text{aq})$ as a reducing agent



Turn over for the next question

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0 7

This question is about ionic solids dissolving in water.

Table 4 shows four enthalpy changes, **A**, **B**, **C** and **D**.

Table 4

	Enthalpy change	$\Delta H / \text{kJ mol}^{-1}$
A	Enthalpy of hydration of chloride ions	−364
B	Enthalpy of hydration of magnesium ions	
C	Enthalpy of lattice dissociation of magnesium chloride	+2493
D	Enthalpy of solution of magnesium chloride	−155

0 7 . 1

Give the meaning of the term enthalpy of hydration.

[1 mark]

The enthalpy change when one mole of gaseous ions dissolves in water to form one mole of aqueous ions.



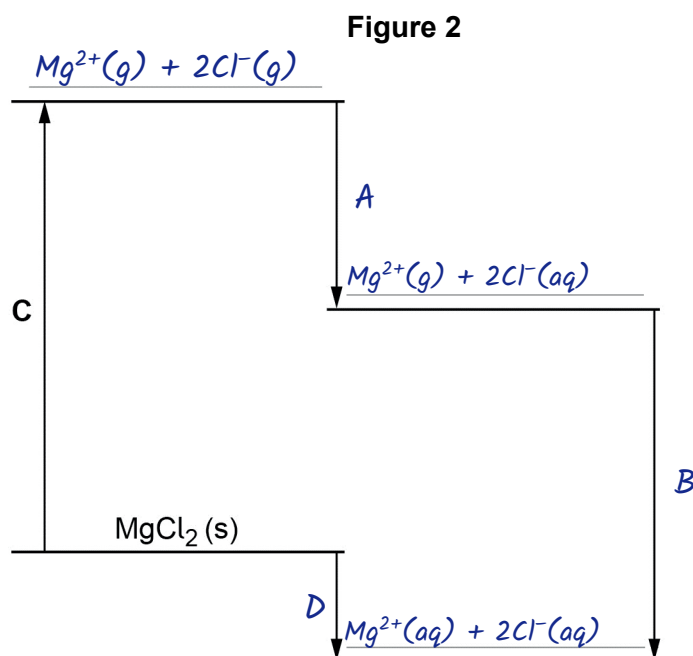
0 7 . 2

Figure 2 shows an incomplete enthalpy cycle for the dissolving of magnesium chloride.

On **Figure 2**

- Write the missing species on each of the three blank lines.
- Label each arrow using the appropriate letters from **Table 4**.

[4 marks]



0 7 . 3

Use the data in **Table 4** to calculate the enthalpy of hydration of magnesium ions.

Show your working.

[2 marks]

$$C + B + 2A = D$$

$$+2493 + B + (2 \times -364) = -155$$

$$B = -155 - 2493 + 728$$

$$B = -1920 \text{ kJ mol}^{-1}$$

Enthalpy of hydration of magnesium ions -1920 kJ mol^{-1}

Turn over ►



07.4

Heat-packs can be used to treat muscular injuries. Some heat-packs work by dissolving an ionic compound in water.

In one type of heat-pack, 10.0 g of magnesium chloride were dissolved in water at 20.0 °C to give 150 g of solution.

Calculate the maximum temperature, in °C, of the heat-pack.
Give your answer to **three** significant figures.

The enthalpy of solution of magnesium chloride = -155 kJ mol^{-1}
The specific heat capacity of the solution = $4.18 \text{ J K}^{-1} \text{ g}^{-1}$

[4 marks]

$$M_r(\text{MgCl}_2) = 24.3 + 71.0 = 95.3$$

$$n(\text{MgCl}_2) = 10.0 \div 95.3 = 0.1049 \text{ mol}$$

$$\text{heat released, } q = 0.1049 \times 155 = 16.26 \text{ kJ} = 16\,264 \text{ J}$$

$$q = mc\Delta T, \text{ so } \Delta T = 16\,264 \div (150 \times 4.18) = 25.9 \text{ K}$$

$$\text{maximum temperature} = 20.0 + 25.9 = 45.9 \text{ }^\circ\text{C}$$

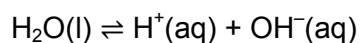
Maximum temperature 45.9 °C

11



0 8

Water is partially ionised.

At 50 °C, $K_w = 5.48 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$

0 8 . 1

Calculate the pH of water at 50 °C

[2 marks]*in pure water $[\text{H}^+] = [\text{OH}^-]$, so $K_w = [\text{H}^+]^2$*

$$[\text{H}^+] = \sqrt{(5.48 \times 10^{-14})} = 2.34 \times 10^{-7} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(2.34 \times 10^{-7}) = 6.63$$

pH 6.63

0 8 . 2

State why water at 50 °C is neutral.

[1 mark]*In water $[\text{H}^+] = [\text{OH}^-]$, so it is neutral even though the pH is below 7.***Question 8 continues on the next page****Turn over ►**

A buffer solution contains 0.100 mol of propanoic acid and 0.200 mol of sodium propanoate in 200 cm³ of solution at 25 °C

0 8 . 3 State what is meant by a buffer solution.

[1 mark]

A solution whose pH remains approximately constant when a small amount of acid or base (alkali) is added.

0 8 . 4 Calculate the pH of the buffer solution at 25 °C
Give your answer to **two** decimal places.

For propanoic acid at 25 °C, the acid dissociation constant, $K_a = 1.35 \times 10^{-5} \text{ mol dm}^{-3}$

[2 marks]

$$K_a = [H^+][A^-] \div [HA], \text{ so } [H^+] = K_a \times [\text{acid}] \div [\text{salt}]$$

$$[H^+] = 1.35 \times 10^{-5} \times 0.500 \div 1.00 = 6.75 \times 10^{-6} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(6.75 \times 10^{-6}) = 5.17$$

pH 5.17

0 8 . 5 20 cm³ of 1.00 mol dm⁻³ aqueous sodium hydroxide were added to the buffer solution.
Calculate the pH of the buffer solution after adding the sodium hydroxide.

Assume that the temperature of the solution stays at 25 °C
Give your answer to **two** decimal places.

[3 marks]

$$n(\text{NaOH}) = 20 \times 1.00 \div 1000 = 0.020 \text{ mol}$$

$$\text{acid left} = 0.100 - 0.020 = 0.080 \text{ mol}$$

$$\text{salt} = 0.200 + 0.020 = 0.220 \text{ mol}$$

$$[H^+] = 1.35 \times 10^{-5} \times (0.080 \div 0.220) = 4.91 \times 10^{-6} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(4.91 \times 10^{-6}) = 5.31$$

pH 5.31



0 9

The equation for the reduction of iron(III) oxide by carbon monoxide is

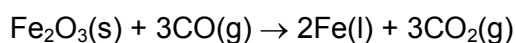


Table 5 gives some entropy values and enthalpies of formation.

Table 5

Substance	$\text{Fe}_2\text{O}_3(\text{s})$	$\text{CO}(\text{g})$	$\text{Fe}(\text{l})$	$\text{CO}_2(\text{g})$
Entropy / $\text{J K}^{-1} \text{mol}^{-1}$	90	198	35	214
Enthalpy of formation / kJ mol^{-1}	-822	-111	+12	-394

0 9

1

Calculate a value for the enthalpy change (ΔH) for this reaction.

[2 marks]

$$\Delta H = \sum \Delta_f H (\text{products}) - \sum \Delta_f H (\text{reactants})$$

$$\Delta H = [(2 \times 12) + (3 \times -394)] - [(-822) + (3 \times -111)]$$

$$\Delta H = -1158 - (-1155) = -3 \text{ kJ mol}^{-1}$$

$$\Delta H \quad \underline{-3} \quad \text{kJ mol}^{-1}$$

0 9

2

Calculate a value for the entropy change (ΔS) for this reaction.

[2 marks]

$$\Delta S = \sum S (\text{products}) - \sum S (\text{reactants})$$

$$\Delta S = [(2 \times 35) + (3 \times 214)] - [90 + (3 \times 198)]$$

$$\Delta S = 712 - 684 = +28 \text{ J K}^{-1} \text{mol}^{-1}$$

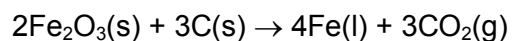
$$\Delta S \quad \underline{+28} \quad \text{J K}^{-1} \text{mol}^{-1}$$

Question 9 continues on the next page

Turn over ►



0 9 . 3 Iron(III) oxide can also be reduced by carbon.



$$\Delta H = +510 \text{ kJ mol}^{-1}$$

$$\Delta S = +585 \text{ J K}^{-1} \text{ mol}^{-1}$$

Explain how the feasibility of this reaction changes as the temperature is increased. You do **not** need to do a calculation.

[2 marks]

$\Delta G = \Delta H - T\Delta S$. As the temperature increases, $T\Delta S$ gets bigger (more positive) because ΔS is positive.

When $T\Delta S$ becomes greater than ΔH ($+510 \text{ kJ mol}^{-1}$), ΔG becomes negative, so the reaction becomes feasible at high temperatures.

6

END OF QUESTIONS



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