

A Level Chemistry A
H432/01 Periodic table, elements and physical chemistry
Sample Question Paper

Date – Morning/Afternoon

Version 2.0

Time allowed: 2 hours 15 minutes

You must have:

- the Data Sheet for Chemistry A

You may use:

- a scientific or graphical calculator



First name

Last name

StemUp Model Solution

Centre
number

Candidate
number

INSTRUCTIONS

- Use black ink. You may use an HB pencil for graphs and diagrams.
- Complete the boxes above with your name, centre number and candidate number.
- Answer **all** the questions.
- Where appropriate, your answers should be supported with working. Marks may be given for a correct method even if the answer is incorrect.
- Write your answer to each question in the space provided.
- Additional paper may be used if required but you must clearly show your candidate number, centre number and question number(s).
- Do **not** write in the bar codes.

INFORMATION

- The total mark for this paper is **100**.
- The marks for each question are shown in brackets [].
- Quality of extended responses will be assessed in questions marked with an asterisk (*).
- This document consists of **28** pages.

SECTION A

You should spend a maximum of 20 minutes on this section.

Answer **all** the questions.

- 1 Which row shows the atomic structure of $^{55}\text{Mn}^{2+}$?

	Protons	Neutrons	Electrons
A	25	30	23
B	25	55	23
C	27	30	25
D	30	25	28

Your answer

A

[1]

- 2 The Group 2 elements react with water, forming a solution and a gas.

Which statement is correct?

- A The reactivity of the elements decreases down Group 2.
 B The pH of the solution formed increases down Group 2.
 C The reaction is a neutralisation.
 D The equation for the reaction of strontium with water is:

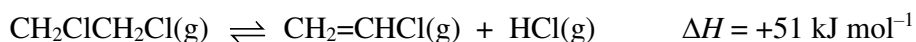


Your answer

B

[1]

- 3 Chloroethene, $\text{CH}_2=\text{CHCl}$, is prepared in the presence of a solid catalyst using the equilibrium reaction below.



Which change would result in an increased equilibrium yield of chloroethene?

- A increasing the pressure
- B increasing the surface area of the catalyst
- C increasing the temperature
- D use of a homogeneous catalyst

Your answer

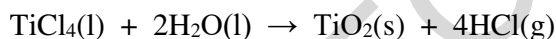
C

[1]

- 4 The table below shows enthalpy changes of formation, $\Delta_f H$.

Compound	$\text{TiCl}_4(\text{l})$	$\text{H}_2\text{O}(\text{l})$	$\text{TiO}_2(\text{s})$	$\text{HCl}(\text{g})$
$\Delta_f H / \text{kJ mol}^{-1}$	-804	-286	-945	-92

What is the value of the enthalpy change of reaction, $\Delta_r H$, for the reaction in the following equation?



- A -63 kJ mol^{-1}
- B -53 kJ mol^{-1}
- C $+53 \text{ kJ mol}^{-1}$
- D $+63 \text{ kJ mol}^{-1}$

Your answer

D

[1]

- 5 Zinc reacts with copper(II) sulfate solution, $\text{CuSO}_4(\text{aq})$.

Which apparatus could be used to determine the effect of the concentration of $\text{CuSO}_4(\text{aq})$ on the rate of reaction?

- A balance
- B gas syringe
- C colorimeter
- D pH meter

Your answer

C

[1]

- 6 The boiling point of hydrogen bromide is $-67\text{ }^\circ\text{C}$.
The boiling point of hydrogen iodide is $-34\text{ }^\circ\text{C}$.

The different boiling points can be explained in terms of the strength of bonds or interactions.

Which bonds or interactions are responsible for the higher boiling point of hydrogen iodide?

- A covalent bonds
- B hydrogen bonds
- C permanent dipole–dipole interactions
- D induced dipole–dipole interactions

Your answer

D

[1]

- 7 The 1st to 8th successive ionisation energies, in kJ mol^{-1} , of an element in period 3 are:

1012 1903 2912 4957 6274 21 269 25 398 29 855

What is the element?

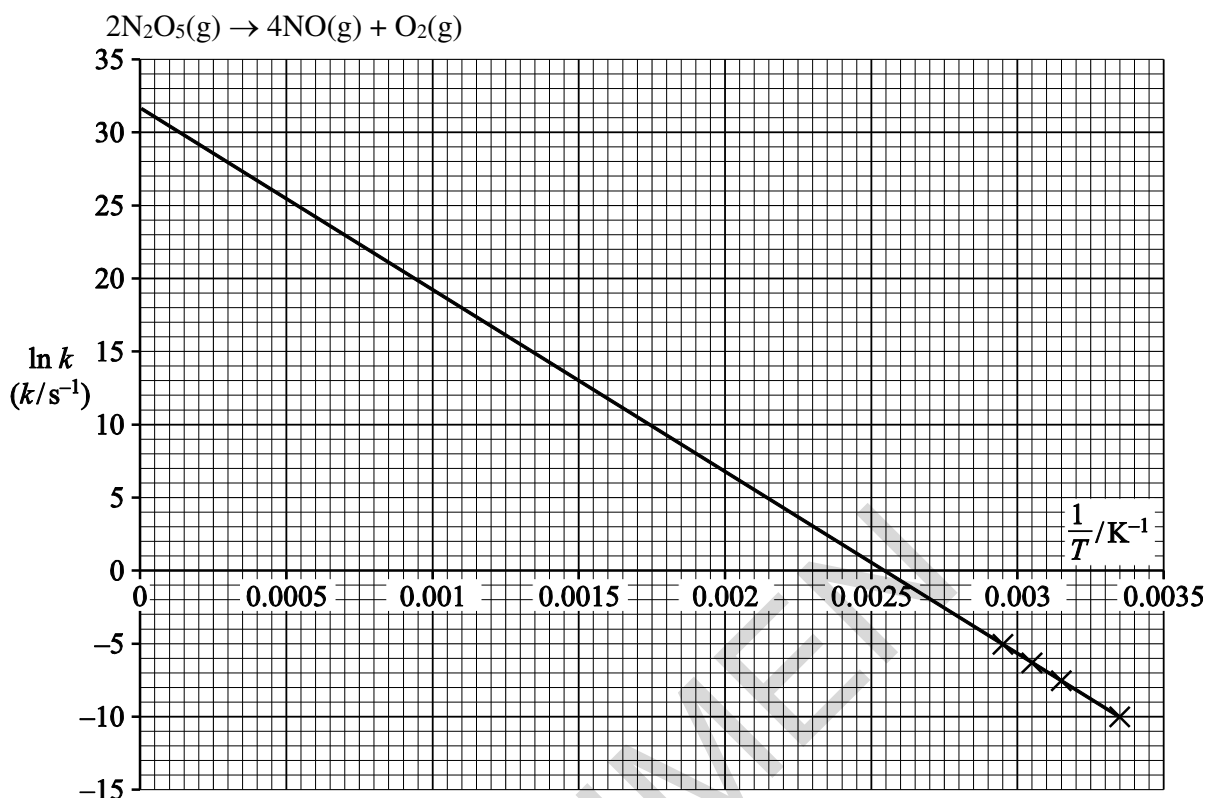
- A Al
- B Si
- C P
- D S

Your answer

C

[1]

- 8 Using the graph, what is the value of the pre-exponential factor, A , for the decomposition of N_2O_5 ?



- A 3.45 s^{-1}
 B 31.5 s^{-1}
 C $1.04 \times 10^5 \text{ s}^{-1}$
 D $4.79 \times 10^{13} \text{ s}^{-1}$

Your answer

D

[1]

- 9 A solution of propanoic acid, $\text{CH}_3\text{CH}_2\text{COOH}$, has a pH of 2.89 at 25°C .

What is $[\text{H}^+]$ in this solution?

- A $1.7 \times 10^{-6} \text{ mol dm}^{-3}$
 B $4.6 \times 10^{-4} \text{ mol dm}^{-3}$
 C $1.3 \times 10^{-3} \text{ mol dm}^{-3}$
 D 0.46 mol dm^{-3}

Your answer

C

[1]

- 10** The lattice enthalpy of calcium chloride can be calculated using **three** of the enthalpy changes below.

Which enthalpy change is **not** required?

- A** enthalpy change of solution of calcium chloride
- B** enthalpy change of hydration of Cl^- ions
- C** enthalpy change of formation of calcium chloride
- D** enthalpy change of hydration of Ca^{2+} ions

Your answer

C

[1]

- 11** Which redox reaction contains the largest change in oxidation state for sulfur?

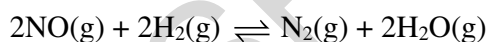
- A** $\text{H}_2\text{SO}_4 + 8\text{HI} \rightarrow \text{H}_2\text{S} + 4\text{I}_2 + 4\text{H}_2\text{O}$
- B** $\text{S} + \text{O}_2 \rightarrow \text{SO}_2$
- C** $\text{S}_2\text{O}_3^{2-} + 2\text{H}^+ \rightarrow \text{SO}_2 + \text{S} + \text{H}_2\text{O}$
- D** $\text{S} + 6\text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + 6\text{NO}_2 + 2\text{H}_2\text{O}$

Your answer

A

[1]

- 12** $\text{NO}(\text{g})$, $\text{H}_2(\text{g})$, $\text{N}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$ exist in equilibrium:



At room temperature and pressure, the equilibrium lies well to the right-hand side.

Which of the following could be the equilibrium constant for this equilibrium?

- A** $1.54 \times 10^{-3} \text{ mol dm}^{-3}$
- B** $6.50 \times 10^2 \text{ mol dm}^{-3}$
- C** $1.54 \times 10^{-3} \text{ dm}^3 \text{ mol}^{-1}$
- D** $6.50 \times 10^2 \text{ dm}^3 \text{ mol}^{-1}$

Your answer

D

[1]

- 13 Copper(II) ions form an aqueous complex ion, **X**, with chloride ions.

Which statement about **X** is true?

- A **X** has optical isomers
- B **X** has a square planar shape
- C **X** has the formula CuCl_4^{2+}
- D **X** has a yellow colour

Your answer

D

[1]

- 14 Two tests are carried out on an aqueous solution of copper(II) sulfate, $\text{CuSO}_4(\text{aq})$.

Test 1: Addition of potassium iodide solution

Test 2: Addition of barium chloride solution

Which of the following statements is/are true?

- 1: **Test 1** produces an off-white precipitate and a brown solution.
- 2: **Test 2** produces a white precipitate.
- 3: **Test 1** and **Test 2** are both redox reactions.

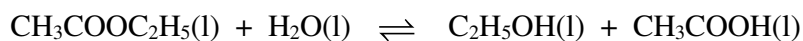
- A 1, 2 and 3
- B Only 1 and 2
- C Only 2 and 3
- D Only 1

Your answer

B

[1]

- 15** Two students set up the equilibrium system below.



The students titrated samples of the equilibrium mixture with sodium hydroxide, $\text{NaOH}(\text{aq})$, to determine the concentration of CH_3COOH .

The students used their results to calculate a value for K_c .

The students' values for K_c were different.

Which of the reason(s) below could explain why the calculated values for K_c were different?

- 1:** Each student carried out their experiment at a different temperature.
- 2:** Each student used a different concentration of $\text{NaOH}(\text{aq})$ in their titration.
- 3:** Each student titrated a different volume of the equilibrium mixture.

- A** 1, 2 and 3
- B** Only 1 and 2
- C** Only 2 and 3
- D** Only 1

Your answer

[1]

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SPECIMEN

SECTION B

Answer **all** the questions.

16 Ammonia is a gas with covalently-bonded molecules consisting of nitrogen and hydrogen atoms.

(a) Show the electron configuration of a nitrogen atom using ‘*electron-in-box*’ diagrams.

Label each sub-shell.



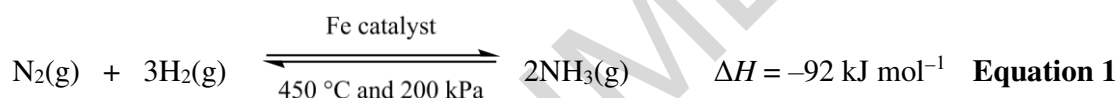
1s

..2s...

..2p...

[2]

(b) Ammonia can be made from the reaction of nitrogen and hydrogen in the Haber process.



What effect will increasing the temperature have on the composition of the equilibrium mixture **and** on the value of the equilibrium constant?

Explain your answer.

..The forward reaction is exothermic, so increasing the temperature favours.....
 ..the backward (endothermic) reaction: the equilibrium shifts to the left.....
 ..There will be more N_2 and H_2 and less NH_3 in the equilibrium mixture,.....
 ..and the value of the equilibrium constant will decrease..... [2]

- (c) A chemist mixes together 0.450 mol N_2 with 0.450 mol H_2 in a sealed container.

The mixture is heated and allowed to reach equilibrium.

At equilibrium, the mixture contains 0.400 mol N_2 and the total pressure is 500 kPa.

Calculate K_p .

Include units in your answer.

At equilibrium: $n(N_2) = 0.400$ mol, so 0.050 mol N_2 reacted

$n(H_2) = 0.450 - 3(0.050) = 0.300$ mol; $n(NH_3) = 2(0.050) = 0.100$ mol

total moles = $0.400 + 0.300 + 0.100 = 0.800$ mol

$p(N_2) = (0.400 \div 0.800) \times 500 = 250$ kPa

$p(H_2) = (0.300 \div 0.800) \times 500 = 187.5$ kPa

$p(NH_3) = (0.100 \div 0.800) \times 500 = 62.5$ kPa

$$K_p = \frac{p(NH_3)^2}{p(N_2) \times p(H_2)^3} = \frac{62.5^2}{250 \times 187.5^3} = 2.37 \times 10^{-6}$$

units: $\frac{kPa^2}{kPa \times kPa^3} = kPa^{-2}$

$K_p = 2.37 \times 10^{-6}$ units kPa^{-2} [5]

- (d) A chemical company receives an order to supply $1.96 \times 10^{10} \text{ dm}^3$ of ammonia at room temperature and pressure. The Haber process produces a 95.0% yield.

Calculate the mass of hydrogen, in tonnes, required to produce the ammonia.

Give your answer to **three** significant figures.

$$n(\text{NH}_3) = 1.96 \times 10^{10} \div 24.0 = 8.167 \times 10^8 \text{ mol}$$

$$n(\text{H}_2) \text{ needed} = 8.167 \times 10^8 \times 3/2 = 1.225 \times 10^9 \text{ mol}$$

$$\text{mass of H}_2 \text{ (100\% yield)} = 1.225 \times 10^9 \times 2.0 = 2.450 \times 10^9 \text{ g}$$

$$= 2.450 \times 10^9 \div 10^6 = 2450 \text{ tonnes}$$

$$\text{for 95.0\% yield: mass} = 2450 \times 100/95.0 = 2579 \text{ tonnes}$$

$$= 2580 \text{ tonnes (3 s.f.)}$$

required mass of hydrogen = ...**2580**..... tonnes [3]

- (e) (i) Hydrazine, N_2H_4 , is used as a rocket fuel. Hydrazine can be prepared from the reaction of ammonia with sodium chlorate(I). There are two other products in the reaction.

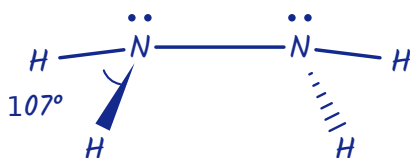
Write an equation for this reaction.



[1]

- (ii) Using the electron pair repulsion theory, draw a 3-D diagram of a molecule of hydrazine.

Predict the H–N–H bond angle around each nitrogen atom.

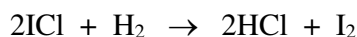


H–N–H bond angle: ..**107°**..... [2]

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SPECIMEN

- 17 Iodine monochloride, ICl, can react with hydrogen to form iodine.



This reaction was carried out several times using different concentrations of ICl or H₂. The initial rate of each experiment was calculated and the results are shown below. Initial concentrations are shown for each experiment.

	[ICl] / mol dm ⁻³	[H ₂] / mol dm ⁻³	Rate / mol dm ⁻³ s ⁻¹
Experiment 1	0.250	0.500	2.04×10^{-2}
Experiment 2	0.500	0.500	4.08×10^{-2}
Experiment 3	0.125	0.250	5.10×10^{-3}

- (a) (i) Calculate the rate constant, k , for this reaction. Include units in your answer.

Exp 1 to 2: [ICl] doubles and the rate doubles, so 1st order in ICl

Exp 1 to 3: [ICl] and [H₂] both halve and the rate falls by a factor of 4 ($2.04 \times 10^{-2} \rightarrow 5.10 \times 10^{-3}$), so 1st order in H₂

$$\text{rate} = k[\text{ICl}][\text{H}_2]$$

$$k = \text{rate} \div ([\text{ICl}][\text{H}_2]) = 2.04 \times 10^{-2} \div (0.250 \times 0.500)$$

$$k = 0.1632 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$$

$$\text{units} = \text{mol dm}^{-3} \text{ s}^{-1} \div (\text{mol dm}^{-3})^2 = \text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$$

$$k = \underline{0.163} \dots \dots \dots \text{units } \underline{\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}} \dots \dots \dots \quad [4]$$

- (ii) Calculate the rate of reaction when ICl has a concentration of $3.00 \times 10^{-3} \text{ mol dm}^{-3}$ and H₂ has a concentration of $2.00 \times 10^{-3} \text{ mol dm}^{-3}$.

$$\text{rate} = k[\text{ICl}][\text{H}_2] = 0.1632 \times 3.00 \times 10^{-3} \times 2.00 \times 10^{-3}$$

$$\text{rate} = \underline{9.79 \times 10^{-7}} \dots \dots \dots \text{mol dm}^{-3} \text{ s}^{-1} \quad [1]$$

- (b) Reaction rates can be increased or decreased by changing the temperature of the reaction. **Fig. 17.1** below shows the energy distribution of the reactant molecules at 25 °C.

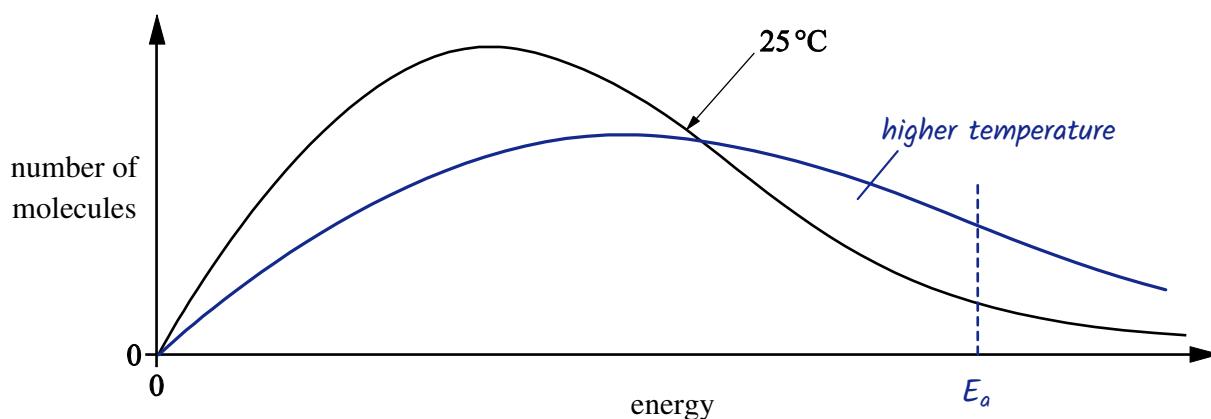


Fig. 17.1

Draw a second curve on **Fig. 17.1**, to represent the distribution of the same number of molecules at a higher temperature.

Use your curve to explain how increasing the temperature increases the rate of reaction.

*At higher temperature the peak is lower and moves right; a greater area.....
 lies beyond E_a , so more molecules have energy greater than or equal to E_a
 So more collisions have enough energy to react, so the rate increases.....*

[2]

18 A student is asked to calculate ΔG at 25 °C for the combustion of butan-1-ol. The teacher provides two pieces of information.

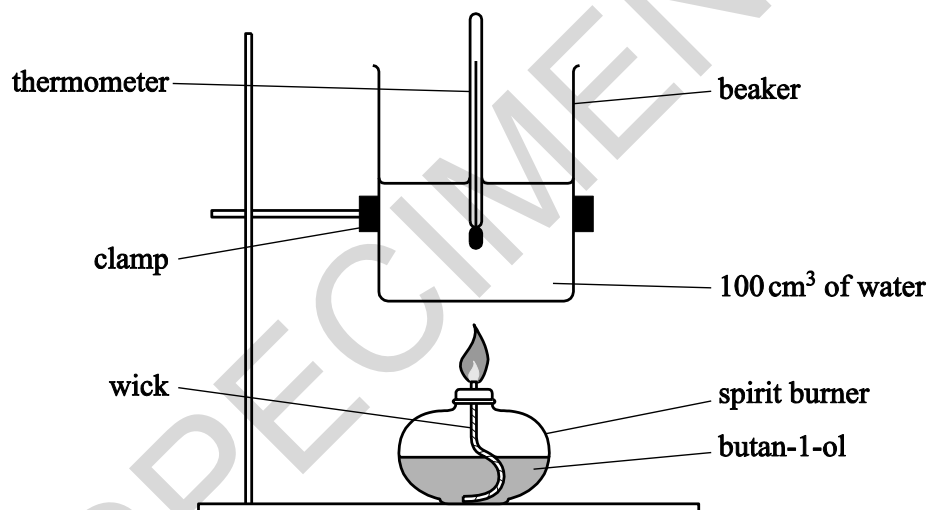
- The equation for the combustion of butan-1-ol.



- Standard entropies of butan-1-ol, oxygen, carbon dioxide and water.

	$\text{CH}_3(\text{CH}_2)_3\text{OH}(\text{l})$	$\text{O}_2(\text{g})$	$\text{CO}_2(\text{g})$	$\text{H}_2\text{O}(\text{l})$
$S^\circ / \text{J K}^{-1} \text{mol}^{-1}$	228	205	214	70

The student carries out an experiment using the apparatus below and obtains the following results. The specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$.



Mass of burner and butan-1-ol before burning / g	98.997
Mass of burner and butan-1-ol after burning / g	98.738
Initial temperature / °C	18.5
Maximum temperature reached / °C	39.0

Use the information on the previous page to calculate ΔG , in kJ mol^{-1} , for the combustion of butan-1-ol according to **Equation 2** at 25°C .

ΔH from the experiment:

$$q = mc\Delta T = 100 \times 4.18 \times (39.0 - 18.5) = 100 \times 4.18 \times 20.5 = 8569 \text{ J} = 8.569 \text{ kJ}$$

$$\text{mass of butan-1-ol burned} = 98.997 - 98.738 = 0.259 \text{ g}$$

$$M(\text{C}_4\text{H}_9\text{OH}) = 74.0 \text{ g mol}^{-1}; \quad n = 0.259 \div 74.0 = 3.50 \times 10^{-3} \text{ mol}$$

$$\Delta H = -8.569 \div 3.50 \times 10^{-3} = -2448 \text{ kJ mol}^{-1}$$

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

$$\Delta S = [(4 \times 214) + (5 \times 70)] - [228 + (6 \times 205)] = 1206 - 1458$$

$$\Delta S = -252 \text{ J K}^{-1} \text{ mol}^{-1} = -0.252 \text{ kJ K}^{-1} \text{ mol}^{-1}$$

$$\Delta G = \Delta H - T\Delta S \quad (T = 298 \text{ K})$$

$$\Delta G = -2448 - (298 \times -0.252) = -2448 + 75.1$$

$$\Delta G = -2373 \text{ kJ mol}^{-1}$$

$$\Delta G = \dots\dots -2373 \dots\dots \text{ kJ mol}^{-1} \quad [7]$$

19 This question is about the chemistry of the elements in Group 2 and the halogens.

- (a) A student prepares an aqueous solution of magnesium chloride by reacting magnesium with excess hydrochloric acid.

Write an equation, including state symbols, for this reaction and state the observation(s) the student should make whilst carrying out this experiment.

equation: $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$

observation(s): *effervescence (fizzing) and the solid Mg dissolves*.....

[2]

- (b) Lattice enthalpies give an indication of the strength of ionic bonding.

How would the lattice enthalpies of magnesium chloride and calcium chloride differ?

Explain your answer.

The lattice enthalpy of MgCl_2 is more exothermic (more negative) than that of CaCl_2

This is because the Mg^{2+} ion is smaller than the Ca^{2+} ion (Mg^{2+} has a greater charge density),.....

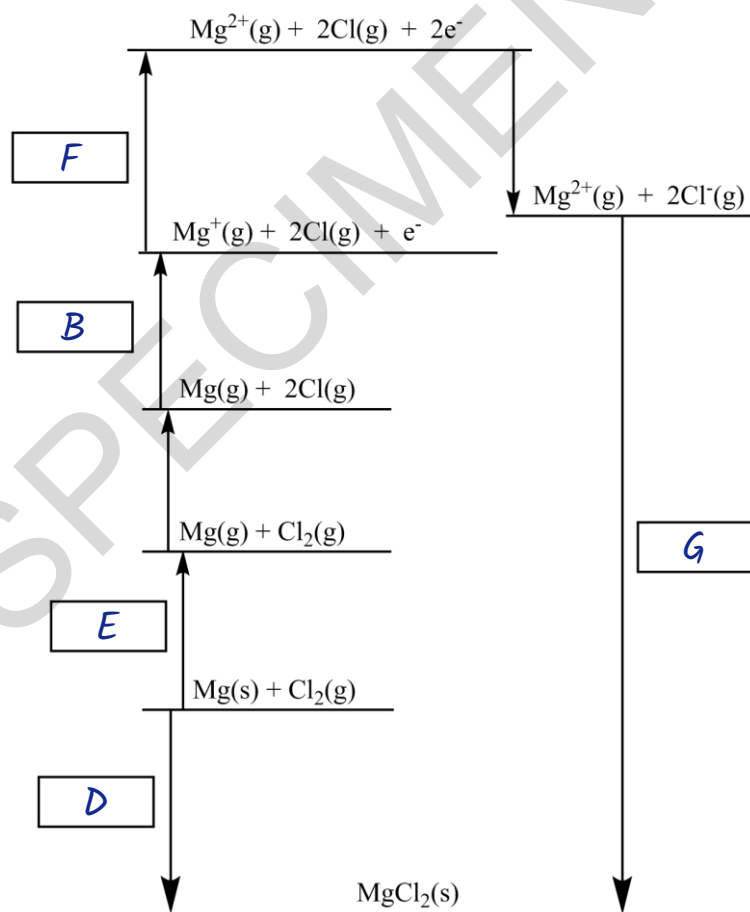
so the attraction between Mg^{2+} and Cl^- ions is greater than between Ca^{2+} and Cl^- ions......

[3]

- (c) The table below shows the enthalpy changes that are needed to determine the lattice enthalpy of magnesium chloride, MgCl_2 .

Letter	Enthalpy change	Energy / kJ mol^{-1}
A	1st electron affinity of chlorine	-349
B	1st ionisation energy of magnesium	+736
C	atomisation of chlorine	+150
D	formation of magnesium chloride	-642
E	atomisation of magnesium	+76
F	2nd ionisation energy of magnesium	+1450
G	lattice enthalpy of magnesium chloride	

- (i) On the cycle below, write the correct letter in each box.



[3]

- (ii) Use the Born-Haber cycle to calculate the lattice enthalpy of magnesium chloride.

$$\Delta H_L = -642 - (+76 + (2 \times 150) + 736 + 1450 + (2 \times -349))$$

$$\Delta H_L = -642 - 1864 = -2506 \text{ kJ mol}^{-1}$$

lattice enthalpy = ...-2506..... kJ mol^{-1} [2]

- (d)* Describe and explain the relative reactivity of the halogens, chlorine, bromine and iodine, in their redox reactions with halides, using reactions on a test-tube scale.

Include reaction equations and observations in your answer.

[6]

Reactivity (oxidising power) decreases down the group: $\text{Cl}_2 > \text{Br}_2 > \text{I}_2$
 Down the group the atoms have more shells and a larger atomic radius,.....
 and there is more shielding by inner electrons. So the attraction of the.....
 nucleus for an incoming electron is weaker, it is harder to gain an electron.....
 and the halogen is a weaker oxidising agent.....
 Test: add chlorine water, bromine water and iodine solution separately to.....
 aqueous KCl, KBr and KI in test tubes, then add cyclohexane and shake to.....
 see the colour of the halogen in the organic layer.....
 Cl_2 with Br^- : the solution turns yellow/orange (orange in cyclohexane).....
 $\text{Cl}_2 + 2\text{Br}^- \rightarrow 2\text{Cl}^- + \text{Br}_2$
 Cl_2 with I^- : the solution turns brown (purple in cyclohexane).....
 $\text{Cl}_2 + 2\text{I}^- \rightarrow 2\text{Cl}^- + \text{I}_2$
 Br_2 with I^- : the solution turns brown (purple in cyclohexane).....
 $\text{Br}_2 + 2\text{I}^- \rightarrow 2\text{Br}^- + \text{I}_2$
 Br_2 with Cl^- and I_2 with Cl^- or Br^- : no reaction, no colour change.....

Additional answer space if required.

So a halogen displaces (oxidises) the halide ions of the halogens below it:.....
 chlorine oxidises Br^- and I^- , bromine oxidises only I^- , and iodine oxidises.....
 neither. The halogen is reduced to halide ions in each reaction. This.....
 confirms the order of reactivity $\text{Cl}_2 > \text{Br}_2 > \text{I}_2$

.....

.....

.....

.....

.....

SPECIMEN

- 20 A student investigates the reactions of two weak monobasic acids: 2-hydroxypropanoic acid, $\text{CH}_3\text{CH}(\text{OH})\text{COOH}$, and butanoic acid, $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$.

- (a) The student wants to prepare a standard solution of 2-hydroxypropanoic acid that has a pH of 2.19.

Plan how the student could prepare 250 cm^3 of this standard solution from solid 2-hydroxypropanoic acid.

In your answer you should provide detail of the practical procedure that would be carried out, including appropriate quantities and necessary calculations.

K_a for 2-hydroxypropanoic acid is $1.38 \times 10^{-4}\text{ mol dm}^{-3}$ at 25°C .

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-2.19} = 6.46 \times 10^{-3}\text{ mol dm}^{-3}$$

$$[\text{CH}_3\text{CH}(\text{OH})\text{COOH}] = [\text{H}^+]^2 \div K_a = (6.46 \times 10^{-3})^2 \div 1.38 \times 10^{-4} \\ = 0.302\text{ mol dm}^{-3}$$

$$n(\text{CH}_3\text{CH}(\text{OH})\text{COOH}) = 0.302 \times 250/1000 = 0.0755\text{ mol}$$

$$M(\text{CH}_3\text{CH}(\text{OH})\text{COOH}) = 90.0\text{ g mol}^{-1}$$

$$\text{mass needed} = 0.0755 \times 90.0 = 6.80\text{ g}$$

..Weigh 6.80 g of the solid accurately and dissolve it in distilled water.....

..(less than 250 cm^3) in a beaker, stirring.....

..Transfer the solution to a 250 cm^3 volumetric flask, rinsing the beaker.....

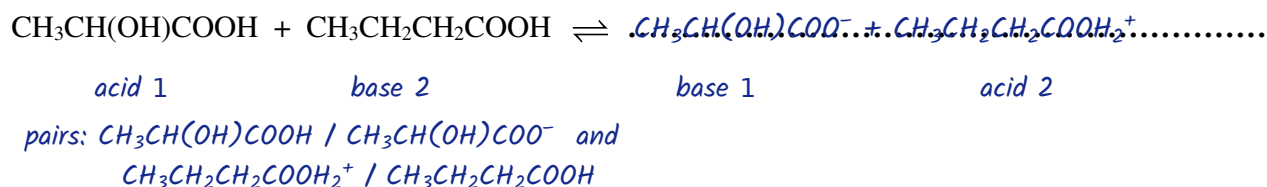
..and rod with distilled water and adding all the washings to the flask.....

..Make the solution up to the 250 cm^3 mark with distilled water and invert.....

..the stoppered flask several times to mix thoroughly..... [8]

- (b) 2-Hydroxypropanoic acid is a slightly stronger acid than butanoic acid. The two acids are mixed together and an acid–base equilibrium is set up.

Suggest the equilibrium equation and identify the conjugate acid–base pairs.



- (c) To prepare a buffer solution, 75.0 cm³ of 0.220 mol dm⁻³ butanoic acid is reacted with 50.0 cm³ of 0.185 mol dm⁻³ sodium hydroxide.

K_a for butanoic acid is 1.5×10^{-5} mol dm⁻³ at 25 °C.

- (i) Calculate the pH of 0.185 mol dm⁻³ sodium hydroxide at 25 °C.

Give your answer to **two** decimal places.

$$[\text{H}^+] = K_w \div [\text{OH}^-] = 1.00 \times 10^{-14} \div 0.185 = 5.41 \times 10^{-14} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(5.41 \times 10^{-14}) = 13.27$$

pH = ...**13.27**..... [2]

- (ii) Calculate the pH of the buffer solution at 25 °C.

Give your answer to **two** decimal places.

$$n(\text{butanoic acid}) \text{ at start} = 0.220 \times 75.0/1000 = 0.0165 \text{ mol}$$

$$n(\text{NaOH}) = 0.185 \times 50.0/1000 = 9.25 \times 10^{-3} \text{ mol} = n(\text{A}^-)$$

$$n(\text{HA}) \text{ left} = 0.0165 - 9.25 \times 10^{-3} = 7.25 \times 10^{-3} \text{ mol}$$

$$[\text{H}^+] = K_a \times [\text{HA}] \div [\text{A}^-] = 1.5 \times 10^{-5} \times (7.25 \times 10^{-3} \div 9.25 \times 10^{-3})$$

$$[\text{H}^+] = 1.18 \times 10^{-5} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(1.18 \times 10^{-5}) = 4.93$$

pH = ...**4.93**..... [4]

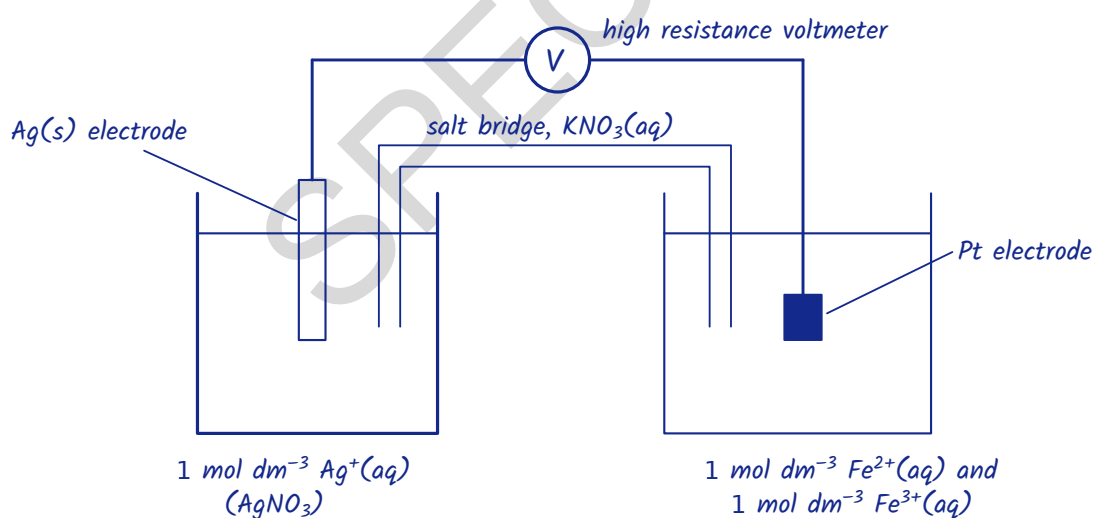
- 21** **Table 21.1** below gives the standard electrode potentials for seven redox systems. You need to use this information to answer the questions below.

Redox system	Equation	E°/V
1	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51
2	$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+1.33
3	$\text{Br}_2(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-(\text{aq})$	+1.09
4	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80
5	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
6	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
7	$\text{Ce}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Ce}(\text{s})$	-2.33

Table 21.1

- (a) (i) Outline an experimental setup that could be used in the laboratory to measure the standard cell potential of an electrochemical cell based on redox systems **4** and **5**.

In your answer you should include details of the apparatus, solutions and the standard conditions required to measure this standard cell potential.



..Standard conditions: 298.K (25.°C) and 100.kPa pressure, with all.....
 ..solutions at 1.mol.dm⁻³ (equal concentrations of Fe²⁺ and Fe³⁺);.....
 ..electrodes joined by wires to a voltmeter, circuit completed by the.....
 ..salt bridge.....

[4]

- (ii) An electrochemical cell can be made based on redox systems **2** and **4**.
The standard cell potential is +0.53 V.

State and explain the effect on the cell potential of this cell if the concentration of silver ions is increased.

The equilibrium $\text{Ag}^+ + \text{e}^- \rightleftharpoons \text{Ag}$ shifts to the right, so the electrode potential of Ag^+/Ag becomes more positive.
Therefore E_{cell} ($1.33 - E(\text{Ag}^+/\text{Ag})$) becomes smaller (less positive). [2]

- (b) From **Table 21.1**, predict the oxidising agent(s) that **will not** oxidise $\text{Fe}^{2+}(\text{aq})$ to $\text{Fe}^{3+}(\text{aq})$.

Zn^{2+} and Ce^{3+} [1]

- (c) An aqueous solution of iron(II) bromide is mixed with an excess of acidified solution containing manganate(VII) ions.

Using **Table 21.1**, give the formulae of the products of any reactions that take place.

Mn^{2+} , H_2O , Fe^{3+} and Br_2
(MnO_4^- , +1.51 V, is more positive than $\text{Fe}^{3+}/\text{Fe}^{2+}$ and Br_2/Br^- , so it oxidises Fe^{2+} to Fe^{3+} and Br^- to Br_2). [2]

- 22** A student carries out a number of experiments on transition metal compounds.

4.800 g of a green hydrated crystalline solid **A** are heated in a crucible to remove the water of crystallisation. 1.944 g of water are removed to leave 0.0180 mol of solid residue **B**.

Solid **B** contains 32.8%, by mass, of the transition metal.

All of **B** is reacted with $\text{AgNO}_3(\text{aq})$ to form 7.695 g of a white precipitate, **C**.

The green crystalline solid **A** is dissolved in water to produce a green solution containing a complex ion, **D**.

When aqueous sodium hydroxide is added to solution of **D**, a grey-green precipitate, **E**, is observed, which dissolves in excess aqueous sodium hydroxide to form a green solution.

- (a) Determine the formulae of **A**, **B**, **D** and **E**.

$$n(\text{AgCl}) = 7.695 \div 143.4 = 0.0537 \text{ mol}$$

$$\text{Cl}^- \text{ per formula of B} = 0.0537 \div 0.0180 = 3, \text{ so B contains 3Cl}$$

$$\text{mass of B} = 4.800 - 1.944 = 2.856 \text{ g}$$

$$M(\text{B}) = 2.856 \div 0.0180 = 158.7 \text{ g mol}^{-1}$$

$$\text{metal: } 32.8\% \times 158.7 = 52.0 \text{ (or } 158.7 - 3 \times 35.5 = 52.2), \text{ so the metal is Cr}$$

$$\text{B} = \text{CrCl}_3 (\text{Cr}^{3+})$$

$$n(\text{H}_2\text{O}) = 1.944 \div 18.0 = 0.108 \text{ mol}$$

$$0.108 \div 0.0180 = 6, \text{ so } 1 \text{ mol CrCl}_3 : 6 \text{ mol H}_2\text{O}$$

D: Cr^{3+} with six water ligands; **E**: chromium(III) hydroxide, which

dissolves in excess NaOH to form green $[\text{Cr}(\text{OH})_6]^{3-}$

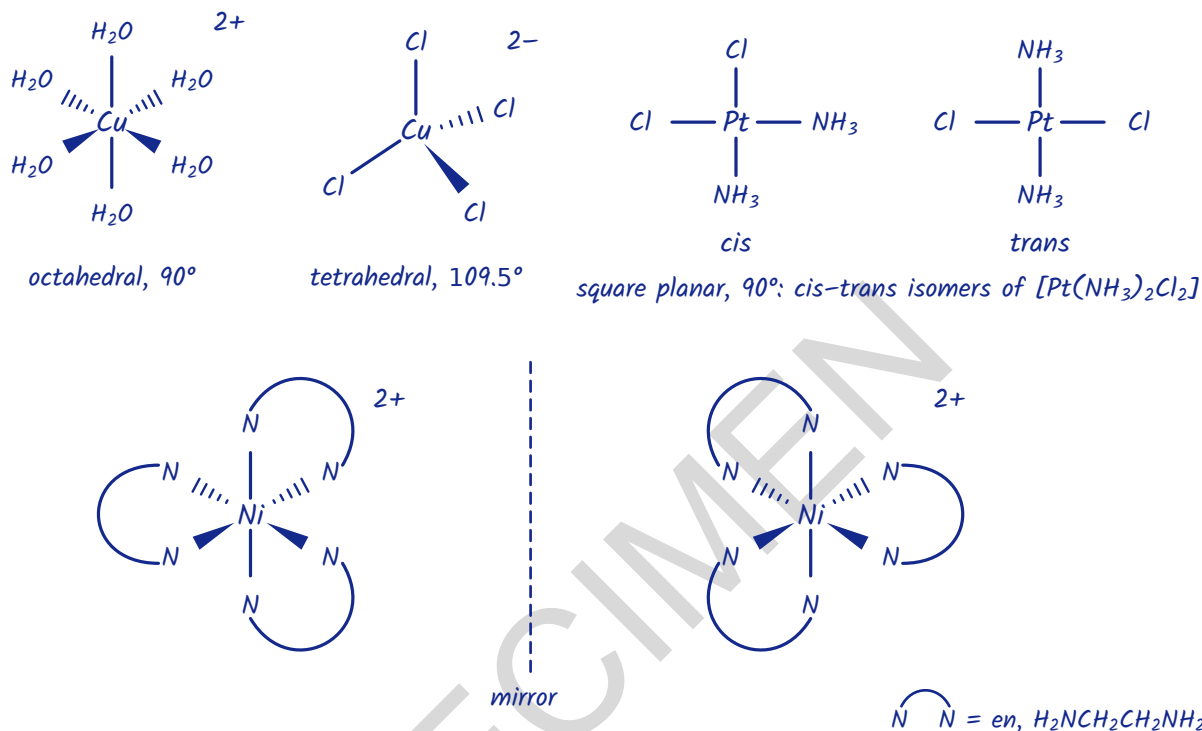


(b)* Transition metal complexes often have different shapes and may form a number of stereoisomers.

Describe the different shapes and the different types of stereoisomerism found in transition metal chemistry.

Use suitable examples and diagrams in your answer.

[6]



optical isomers of $[\text{Ni}(\text{en})_3]^{2+}$: non-superimposable mirror images

Shapes: complexes with six coordinate bonds (six monodentate ligands, e.g. $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$) are octahedral with 90° bond angles.

Complexes with four coordinate bonds are tetrahedral (e.g. $[\text{CuCl}_4]^{2-}$, 109.5° , with large Cl^- ligands) or square planar (e.g. $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$, 90°).

Cis-trans isomerism occurs in square planar and octahedral complexes with two different types of ligand. In *cis*- $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$ (cisplatin, an anticancer drug) the two Cl ligands are adjacent (same side, 90° apart); in the *trans* isomer they are opposite (180°). Octahedral $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]^+$ also has *cis* (Cl adjacent) and *trans* (Cl opposite) isomers.

Optical isomerism occurs in octahedral complexes with bidentate ligands.

Additional answer space if required.

such as en ($\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2$), e.g. $[\text{Ni}(\text{en})_2]^{2+}$ or cis- $[\text{Co}(\text{en})_2\text{Cl}_2]^+$. The two optical isomers are non-superimposable mirror images of each other (as drawn above), so they cannot be overlaid. Cis- $[\text{Co}(\text{en})_2\text{Cl}_2]^+$ shows both types: it is a cis-trans isomer of the trans form and also exists as a pair of optical isomers.

END OF QUESTION PAPER

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