A Level Chemistry A

H432/01 Periodic table, elements and physical chemistry

Sample Question Paper

Date – Morning/Afternoon

Time allowed: 2 hours 15 minutes

You must have:
- the Data Sheet for Chemistry A

You may use:
- a scientific calculator

First name									
Last name									
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.
- 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 $^{55}Mn^{2+}$?

→ from periodic table

atomic no. = 25 = # protons

	Protons	Neutrons	Electrons
[✓ **A**]	25	30	23
B	25	55	23
~~**C**~~	27	30	25
~~**D**~~	30	25	28

neutrons: atomic mass − # protons = 55 − 25 = 30

Your answer A

$\# e^- = 25 - 2 = 23$ (as +2 charge)

[1]

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

Which statement is correct?

(increase)

~~**A**~~ The reactivity of the elements decreases down Group 2.

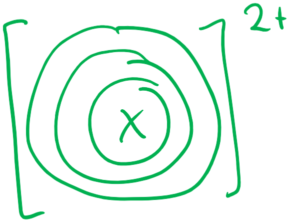
[✓ **B**] The pH of the solution formed increases down Group 2.

~~**C**~~ The reaction is a neutralisation. (redox)

~~**D**~~ The equation for the reaction of strontium with water is:

$$2Sr + 2H_2O \rightarrow 2SrOH + H_2$$

→ ion is less charge dense down the group ∴ ionic bonding in X(OH$_2$) is weaker so OH$^-$ more readily released + pH ↑.



↳ correct equation: $X + 2H_2O \rightarrow X(OH)_2 + H_2$

Your answer B

[1]

→ favours reverse reaction as produces fewer moles

3 Chloroethene, $CH_2{=}CHCl$, is prepared in the presence of a solid catalyst using the equilibrium reaction below.

heterogeneous

$$CH_2ClCH_2Cl(g) \rightleftharpoons CH_2{=}CHCl(g) + HCl(g) \qquad \Delta H = +51\ \text{kJ mol}^{-1}$$

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

le Chatelier's principle

~~**A**~~ increasing the pressure

~~**B**~~ increasing the surface area of the catalyst → ↑ ROR for forward + reverse reaction

[✓ **C**] increasing the temperature → as forward reaction is endothermic

~~**D**~~ use of a homogeneous catalyst → same effect on forward + reverse reaction

Your answer C

[1]

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

Compound	$TiCl_4(l)$	$H_2O(l)$	$TiO_2(s)$	$HCl(g)$
$\Delta_f H$ / kJ mol^{-1}	–804	–286	–945	–92

Hess cycle

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

$$TiCl_4(l) + 2H_2O(l) \rightarrow TiO_2(s) + 4HCl(g)$$

A –63 kJ mol^{-1}

B –53 kJ mol^{-1}

C +53 kJ mol^{-1}

[✓ **D**] +63 kJ mol^{-1}

Your answer D

[1]

$$TiCl_4 + 2H_2O \xrightarrow{\Delta H_1} TiO_2 + 4HCl.$$

$\Delta H_2 \uparrow \qquad \uparrow \Delta H_3$

$Ti + H_2 + O_2 + Cl_2$

CW = ACW

$\Delta H_1 + \Delta H_2 = \Delta H_3$

$\Delta H_1 = \Delta H_3 - \Delta H_2$

$\Delta H_1 = [-945 + (4 \times -92)] - [-804 + (2 \times -286)]$

$\Delta H_1 = -1313 - (-1376) = \underline{+63\ \text{kJmol}^{-1}}$

not a transition metal

→ transition metal compound ∴ coloured.

5 Zinc reacts with copper(II) sulfate solution, $CuSO_4(aq)$.

Which apparatus could be used to determine the effect of the concentration of $CuSO_4(aq)$ on the rate of reaction?

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

Blue sol. colourless sol.

~~**A**~~ balance

~~**B**~~ gas syringe

[✓ **C** colorimeter]

~~**D**~~ pH meter

Displacement reaction: Zn is more reactive than Cu.

Your answer ☐

[1]

H—Br

H—I

6 The boiling point of hydrogen bromide is –67 °C.
The boiling point of hydrogen iodide is –34 °C.

HBr
HI]→ simple molecular structure

→ Iodine is not electronegative enough to form hydrogen bonds + is less electronegative than bromine.

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]

→ as iodine has more e⁻ than bromine

covalent bonds

H—I

intermolecular forces → broken when we boil HI or HBr

H—I

H—I

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

big 'jump' indicates moving to an inner shell.

→ 5 e⁻ in outer shell so element is in group 5 (15) ∴ P

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?

$2N_2O_5(g) \rightarrow 4NO(g) + O_2(g)$

A 3.45 s^{-1}

B 31.5 s^{-1}

C 1.04×10^{5} s^{-1}

[✓ **D**] 4.79×10^{13} s^{-1}

Arrhenius equation: (linear form)

$\ln k = \frac{-E_A}{RT} + \ln A$

$y = mx + c$

y-intercept = 31.5 = ln A

$A = e^{31.5} = 4.79 \times 10^{13}\ s^{-1}$

Your answer D

[1]

9 A solution of propanoic acid, CH_3CH_2COOH, has a pH of 2.89 at 25 °C.

What is $[H^+]$ in this solution?

A 1.7×10^{-6} mol dm^{-3}

B 4.6×10^{-4} mol dm^{-3}

[✓ **C**] 1.3×10^{-3} mol dm^{-3}

D 0.46 mol dm^{-3}

$pH = -\log_{10}[H^+]$

$[H^+] = 10^{-pH} = 10^{-2.89}$

$[H^+] = 1.288... \times 10^{-3}$ $moldm^{-3}$

$\therefore 1.3 \times 10^{-3}$ $moldm^{-3}$

Your answer C

[1]

→ forming solid, ionic compound from the respective gaseous ions → ΔH_{LE}

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

using Born-Haber cycles.

Which enthalpy change is **not** required?

$\Delta H_{sol.}$ **A** enthalpy change of solution of calcium chloride

$\Delta H_{hyd.}$ **B** enthalpy change of hydration of Cl^- ions

ΔH_f **C** enthalpy change of formation of calcium chloride

$\Delta H_{hyd.}$ **D** enthalpy change of hydration of Ca^{2+} ions

$Ca^{2+}(g) + 2Cl^-(g)$

↓ ΔH_{LE}

$CaCl_2(s)$

↓ $\Delta H_{sol.}$

$CaCl_2(aq)$

$\Delta H_{hyd.}$ $(Ca^{2+} + 2Cl^-)$

Used for a larger Born-Haber cycle.

Your answer C

[1]

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

[✓ **A**] $\overset{+2\,+6\,-8}{H_2SO_4} + 8HI \rightarrow \overset{+2\,-2}{H_2S} + 4I_2 + 4H_2O$ (−8)

B $\overset{0}{S} + O_2 \rightarrow \overset{+4\,-4}{SO_2}$ (+4)

C $\overset{+2\,-6}{S_2O_3^{2-}} + 2H^+ \rightarrow \overset{+4}{SO_2} + \overset{0}{S} + H_2O$ (2)

D $\overset{0}{S} + 6HNO_3 \rightarrow \overset{+6}{H_2SO_4} + 6NO_2 + 2H_2O$ (+6)

– = ox. state of sulfur

– = ox. state of other elements.

$O_4 \rightarrow O = -2$

$-2 \times 4 = -8.$

Your answer A

[1]

12 $NO(g)$, $H_2(g)$, $N_2(g)$ and $H_2O(g)$ exist in equilibrium:

$$2NO(g) + 2H_2(g) \rightleftharpoons N_2(g) + 2H_2O(g)$$

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 1.54×10^{-3} mol dm^{-3}

★ **B** 6.50×10^{2} mol dm^{-3}

C 1.54×10^{-3} dm^3 mol^{-1}

★ [✓ **D**] 6.50×10^{2} dm^3 mol^{-1}

$$K_c = \frac{[N_2][H_2O]^2}{[NO]^2[H_2]^2}$$

$K_c < 1 \leftarrow 1 \rightarrow K_c => 1$

Your answer D

[1]

$$\text{units} = \frac{\cancel{mol\,dm^{-3}} \times \cancel{mol\,dm^{-3}} \times \cancel{mol\,dm^{-3}}}{\cancel{mol\,dm^{-3}} \times \cancel{mol\,dm^{-3}} \times \cancel{mol\,dm^{-3}} \times mol\,dm^{-3}} = \frac{1}{mol\,dm^{-3}}$$

$= mol^{-1}\,dm^3$

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

RECALL

Which statement about **X** is true?

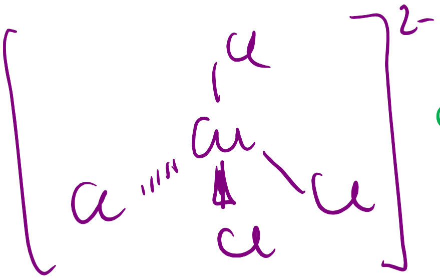
$[CuCl_4]^{2-}$ tetrahedral

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



← 4 of the same ligand bonded to Cu^{2+} so no optical isomers.

Your answer D

[1]

14 Two tests are carried out on an aqueous solution of copper(II) sulfate, $CuSO_4(aq)$.

Test 1: Addition of potassium iodide solution
Test 2: Addition of barium chloride solution → test for SO_4^{2-} → ✓ white ppt. ($BaSO_4(s)$)

Which of the following statements is/are true?

✓ **1:** **Test 1** produces an off-white precipitate and a brown solution.

$CuI(s)$ $I_2(aq)$

✓ **2:** **Test 2** produces a white precipitate.

off-white ppt. ↑ brown sol. ↑

✗ **3:** **Test 1** ✓ and **Test 2** ✗ are both redox reactions.

① redox: $2Cu^{2+}(aq) + 4I^-(aq) \rightarrow 2CuI(s) + I_2(aq)$

A 1, 2 and 3

[✓ **B**] Only 1 and 2

C Only 2 and 3

D Only 1

(white ppt.)

② $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

↳ no change in ox. states ∴ not redox.

Your answer B

[1]

15 Two students set up the equilibrium system below.

$$CH_3COOC_2H_5(l) + H_2O(l) \rightleftharpoons C_2H_5OH(l) + CH_3COOH(l)$$

The students titrated samples of the equilibrium mixture with sodium hydroxide, NaOH(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?

SAME RATIOS ⇑

✓ **1:** Each student carried out their experiment at a different temperature.

✗ **2:** Each student used a different concentration of NaOH(aq) in their titration. ⇒ would need ↑/↓ volume of NaOH so same moles would react.

✗ **3:** Each student titrated a different volume of the equilibrium mixture.

Kc is only affected by changes in temperature (affects ⇌ ratios that don't restore).

~~**A**~~ 1, 2 and 3

~~**B**~~ Only 1 and 2

~~**C**~~ Only 2 and 3

[✓ **D**] Only 1

Your answer D

[1]

⇒ would need ↑ / ↓ volume of NaOH but same moles would react ⇒ SAME RATIOS

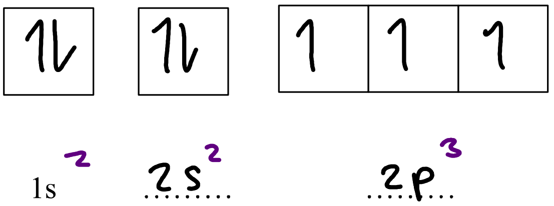
SECTION B

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.



Hund's Rule
fill each orbital
singularly before
filling an orbital
with 2 electrons

[2]

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

$$N_2(g) + 3H_2(g) \underset{450\ °C\ \text{and}\ 200\ kPa}{\overset{\text{Fe catalyst}}{\rightleftharpoons}} 2NH_3(g) \qquad \Delta H = -92\ kJ\ mol^{-1}$$ **Equation 1**

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 temperature favours the backwards reaction so more N_2 and H_2 produced so 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.

Show **all** your working.

Include units in your answer.

	N_2	H_2	NH_3
I	0.450	0.450	0
C	−0.050	−0.150	+0.100
E	0.400	0.300	0.100

$$p(N_2) = \frac{0.400}{0.800} \times 500 = 250\ kPa$$

$$p(H_2) = \frac{0.300}{0.800} \times 500 = 187.5\ kPa$$

$$p(NH_3) = \frac{0.100}{0.800} \times 500 = 62.5\ kPa$$

Partial pressure = mol fraction × total pressure

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

$$\frac{\cancel{kPa^2}}{\cancel{kPa} \times kPa^{\cancel{3}2}} = \frac{1}{kPa^2}$$

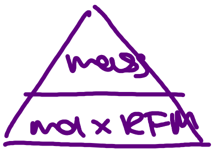
K_p = 2.37×10^{-6} units kPa^{-2} **[5]**

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

$n = \frac{\text{volume (dm}^3\text{)}}{24}$

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

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



$n(NH_3) = \frac{1.96 \times 10^{10}}{24} = 8.167 \times 10^8$ mol

$n(H_2) = \frac{8.167 \times 10^8}{2} \times 3 = 1.225 \times 10^9$ mol

$2 \times 1.225 \times 10^9 = 2.450 \times 10^9$ g

mass of $H_2 = \frac{2.450 \times 10^9}{1 \times 10^6} = 2450$ tonnes

$\frac{2450 \times 100}{95} = 2580$ tonnes

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.

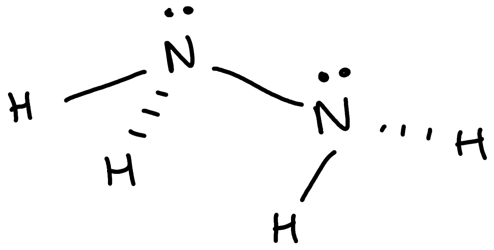
$2NH_3 + NaOCl \longrightarrow N_2H_4 + NaCl + H_2O$

[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.

3bp : 1lp pyramidal 107°



H–N–H bond angle: 107° **[2]**

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

$$2ICl + H_2 \rightarrow 2HCl + I_2$$

This reaction was carried out several times using different concentrations of ICl or H_2. 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 ×2	0.500 ×1	4.08×10^{-2} ×2
Experiment 3	0.125 ÷4	0.250 ÷2	5.10×10^{-3} ÷8

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

Show **all** your working.

$rate = k[ICl]^x[H_2]^y$

x and y indicate orders of reaction

$[H_2]$ and $[ICl]$ are both first order reactions as their increase in concentration is mirrored by the change in rate of reaction.

$rate = k[ICl][H_2]$

$$\frac{2.04 \times 10^{-2}}{0.250 \times 0.500} = 0.163\ mol^{-1}\,dm^{3}\,s^{-1}$$

$$\frac{\cancel{mol\,dm^{-3}}\ s^{-1}}{mol^{2}\,dm^{-6}}$$

k = 0.163 units $mol^{-1}\,dm^{3}\,s^{-1}$ **[4]**

(ii) Calculate the rate of reaction when ICl has a concentration of 3.00×10^{-3} mol dm^{-3} and H_2 has a concentration of 2.00×10^{-3} mol dm^{-3}.

Show **all** your working.

$rate = 0.163[ICl][H_2]$

$rate = 0.163[3 \times 10^{-3}][2 \times 10^{-3}]$

$rate = 9.78 \times 10^{-7}$ $mol\,dm^{-3}\,s^{-1}$

rate = 9.78×10^{-7} mol dm^{-3} s^{-1} **[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.

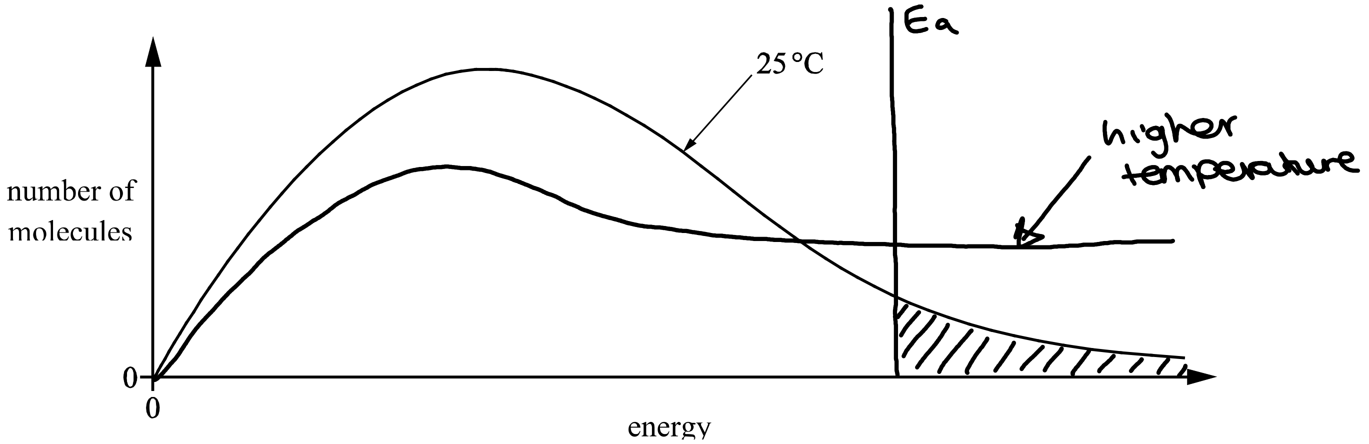


Fig. 17.1

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.

The graph shows that at a higher temperature more molecules have energy $> E_a$

[2]

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

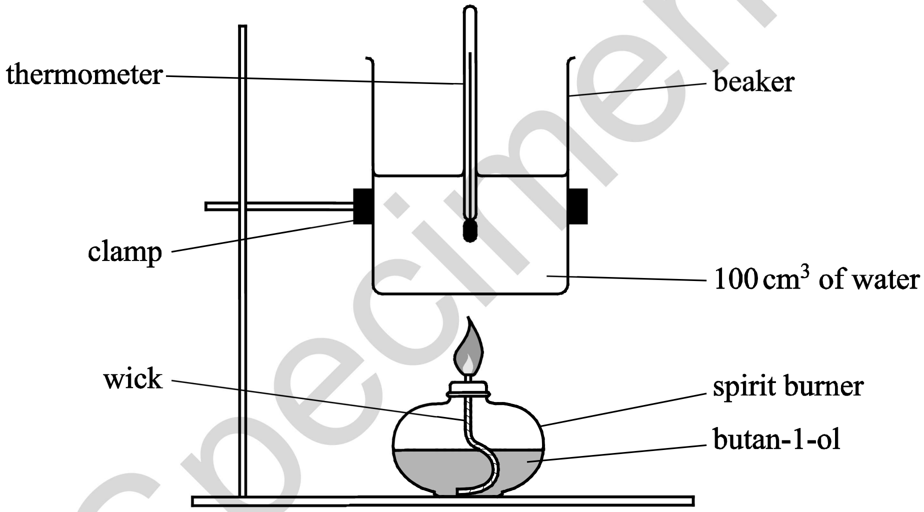
- The equation for the combustion of butan-1-ol.

$CH_3(CH_2)_3OH(l) \ + \ 6O_2(g) \ \rightarrow \ 4CO_2(g) \ + \ 5H_2O(l)$ **Equation 2**

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

	$CH_3(CH_2)_3OH(l)$	$O_2(g)$	$CO_2(g)$	$H_2O(l)$
S°/ J K^{-1} mol^{-1}	228	205	214	70

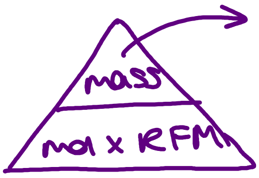
The student carries out an experiment using the apparatus below and obtains the following results. The specific heat capacity of water is 4.18 J g^{-1} K^{-1}.



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 / ºC	18.5
Maximum temperature reached / ºC	39.0

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

Show **all** your working.



98.997 − 98.738 = 0.259 g of butan-1-ol

energy → $q = m\,c\,\Delta T$

mass of water

specific heat capacity

temperature change

$\Delta G = \Delta H - T\Delta S$

$q = 100 \times 4.18 \times (39 - 18.5) = 8569\,J \rightarrow 8.569\,kJ$

moles of butan-1-ol : $\frac{0.259}{74} = 3.5 \times 10^{-3}$ mol

$\Delta S = \Delta S_{products} - \Delta S_{reactants}$

$\Delta S = (4 \times 214) + (5 \times 70) - [(228) + (6 \times 205)] = -252\,J\,K^{-1}\,mol^{-1}$

$\Delta S = -0.252\,kJ\,K^{-1}\,mol^{-1}$

$\Delta H = \frac{8.569}{3.5 \times 10^{-3}} = -2448\,kJ\,mol^{-1}$

$\Delta G = -2448 - (298 \times -0.252) = -2373\,kJ\,mol^{-1}$

Δ*G* = -2373 kJ mol^{-1} **[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.

HCl $MgCl_2$ Mg

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

equation: $Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$

observation(s): effervescence and solid 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.

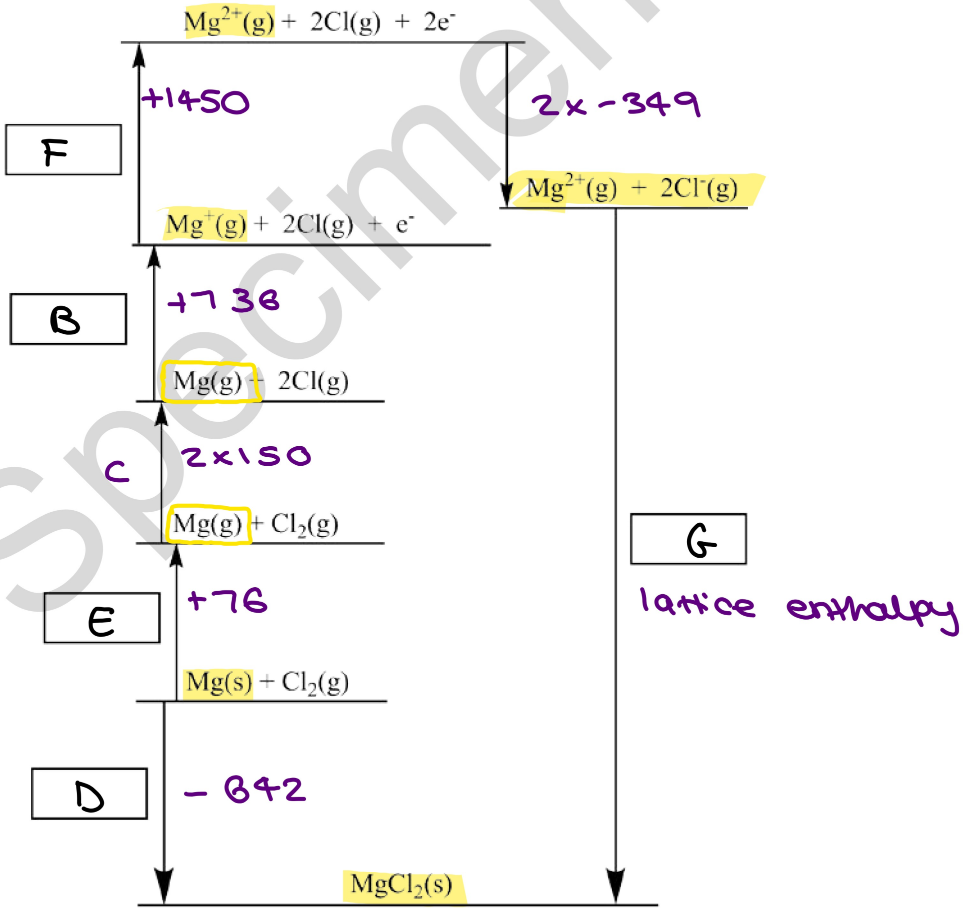
lattice enthalpy of $MgCl_2$ is more exothermic than $CaCl_2$ because Mg^{2+} is smaller than Ca^{2+} therefore, the attraction between Mg^{2+} and Cl^- is greater

[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.

$-(2x-349) - 1450 - 736 - (2\times150) - 76 - 642$

$= -2506$ kJmol^{-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.

$Cl_2 + 2Br^- \longrightarrow Br_2 + 2Cl^-$ yellow solution

$Cl_2 + 2I^- \longrightarrow I_2 + 2Cl^-$ orange / brown solution

Down the group:

- more shells
- increased shielding
- more difficult to gain an electron

[6]

20 A student investigates the reactions of two weak monobasic acids: 2-hydroxypropanoic acid, $CH_3CH(OH)COOH$, and butanoic acid, $CH_3CH_2CH_2COOH$.

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

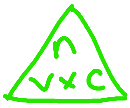
Plan how the student could prepare 250 cm^3 of this standard solution from solid 2-hydroxypropanoic acid.

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

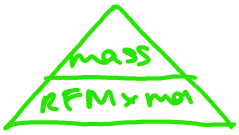
K_a for 2-hydroxypropanoic acid is 1.38×10^{-4} mol dm^{-3} at 25 °C.

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

$[CH_3CH(OH)COOH] = \frac{[H^+]^2}{Ka} = \frac{[6.46 \times 10^{-3}]^2}{1.38 \times 10^{-4}} = 0.0302$ moldm^{-3}



$0.0302 \times 250 \times 10^{-3} = 0.0755$ mol of 2-hydroxypropanoic acid



$0.0755 \times 90 = 6.80$g of 2-hydroxypropanoic acid

$\hookrightarrow 12+3+12+1+16+1+12+(16\times2)+1$

- Dissolve 6.80g of solid in less than 250 cm³ distilled water in a beaker
- transfer the solution to a 250 cm³ volumetric flask and ensure all solution is washed out of the beaker
- make solution up to 250 cm³ with distilled water and ensure thorough mixing by inverting.

[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.

$CH_3CH(OH)COOH + CH_3CH_2CH_2COOH \rightleftharpoons$ $CH_3CH(OH)COO^- + CH_3CH_2CH_2COOH_2^+$

acid 1   base 2   base 1   acid 2

[2]

(c) To prepare a buffer solution, 75.0 cm^3 of 0.220 mol dm^{-3} butanoic acid is reacted with 50.0 cm^3 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.

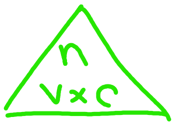
$pOH + pH = 14$ $pOH = -\log_{10}[OH^-]$

$pOH = -\log_{10}(0.185) = 0.73$

$14 - 0.73 = 13.27$

pH = 13.27 **[2]**

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



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

Show **all** your working.

NaOH: $50\times10^{-3} \times 0.185 = 9.25\times10^{-3}$ mol

butanoic acid: $75\times10^{-3} \times 0.22 = 0.0165$ mol

$0.0165 - 9.25\times10^{-3} = 7.25\times10^{-3}$ mol of buffer

NaOH: $9.25\times10^{-3} \div 125\times10^{-3} = 0.074$ moldm^{-3}

buffer: $7.25\times10^{-3} \div 125\times10^{-3} = 0.058$ moldm^{-3}

$Ka = \frac{[H^+][A^-]}{[HA]}$ $[H^+] = 1.5\times10^{-5} \times \frac{0.058}{0.074} = 1.176$ moldm^{-3}

$pH = -\log_{10}(1.176) = 4.93$

$[H^+] = Ka \times \frac{[HA]}{[A^-]}$ $pH = -\log_{10}[H^+]$

pH = 4.93 **[4]**

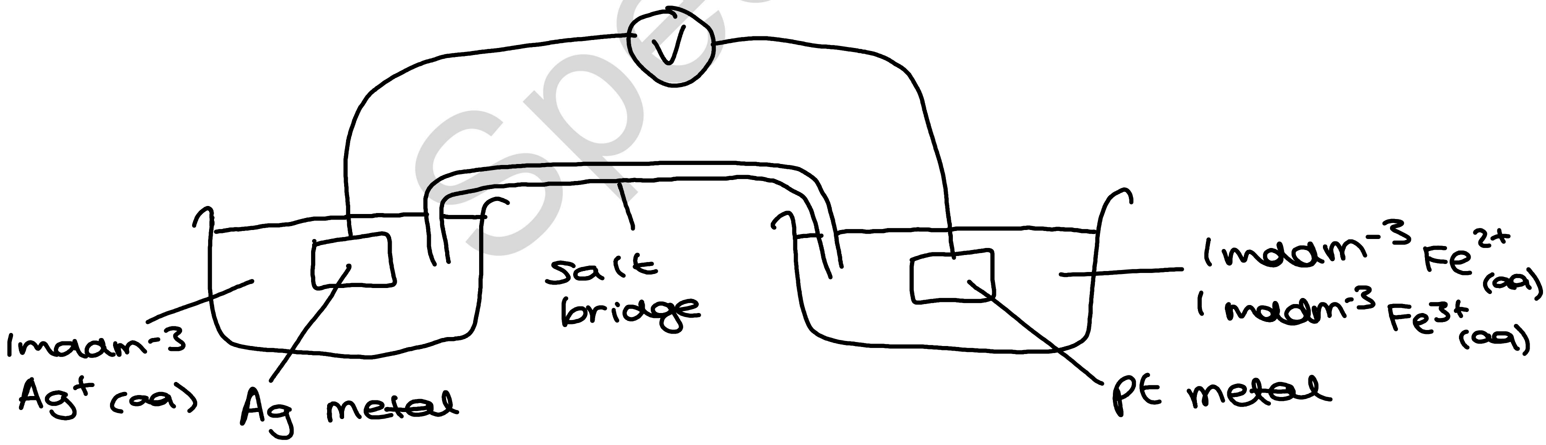
21 **Table 21.1** below gives the standard electrode potentials for seven redox systems. You need to use this information to answer the questions below.

Redox system	Equation	$E^\ominus$/V
1	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
2	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
3	$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
4	$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
5	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$ no solid	+0.77
6	$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	–0.76
7	$Ce^{3+}(aq) + 3e^- \rightleftharpoons Ce(s)$	–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: 298K, 100kPa ....

....

....

.... **[4]**

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

redox system 4 shifts to the right

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

Ag/Ag^+ becomes more positive so, E_{cell} becomes less positive

[2]

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

Ce^{3+}, Zn^{2+}

$E^\ominus$ will need to be less than +0.77 V

[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+}, 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 $AgNO_3(aq)$ to form 7.695 g of a white precipitate, **C**.

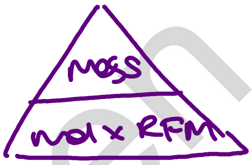
→ test for Cl^- ions

AgCl

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**.



Show **all** your working.

molar ratio

$$\frac{7.695}{108+35.5} = 0.05362 \text{ mol of AgCl}$$

$$\frac{0.05362}{0.0180} = 3 \; Cl^- \text{ ions}$$

$$\frac{4.800-1.944}{0.0180} = 158.7 \text{ gmol}^{-1} \; B$$

$$158.7 - (35.5 \times 3) = 52.2 = Cr$$

$$\frac{1.944}{16+2} = 0.108 \text{ mol of } H_2O$$

$$\frac{0.108}{0.0180} = 6 \text{ mol of } H_2O$$

A = $CrCl_3 \cdot 6H_2O$

D = $[Cr(H_2O)_6]^{3+}$

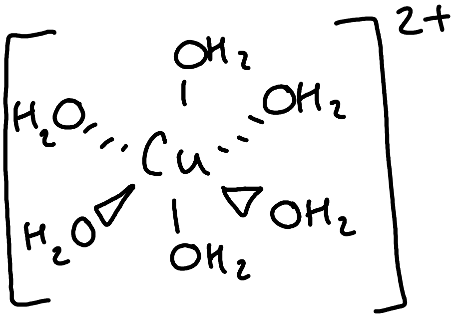
B = $CrCl_3$

E = $Cr(OH)_3$ **[9]**

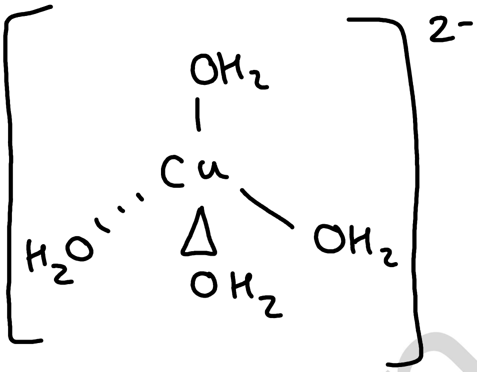
(b)* Transition metal complexes often have different shapes and may form a number of stereoisomers.

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

Use suitable examples and diagrams in your answer.

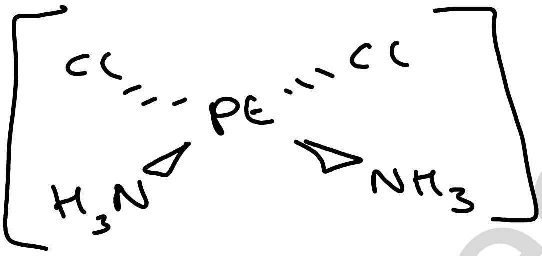


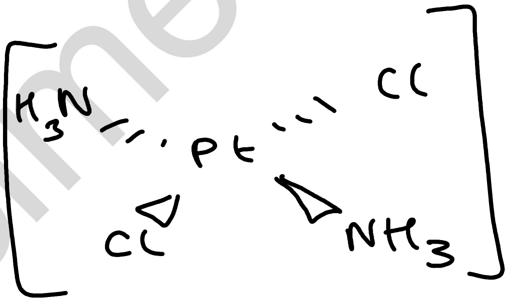
octahedral



tetrahedral

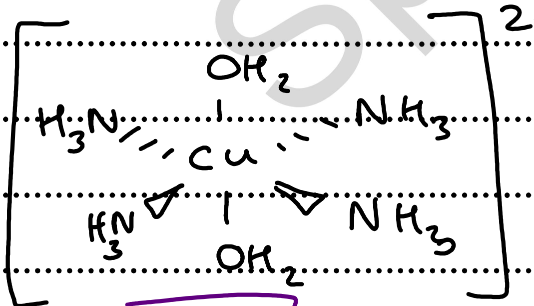
cis



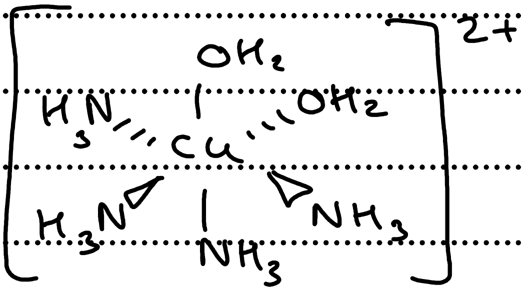


trans

square planar

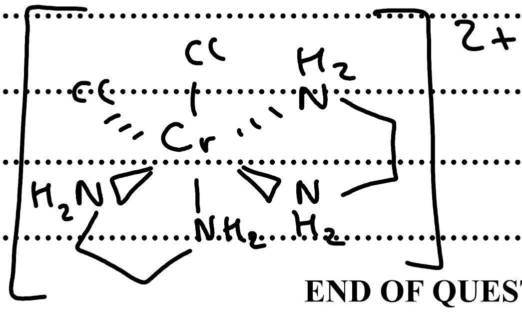


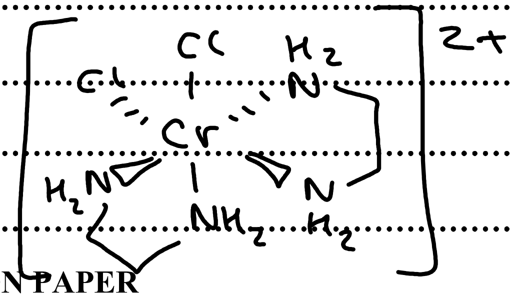
trans



cis

OPTICAL: bidentate ligands





[6]

END OF QUESTION PAPER