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Pearson Edexcel Level 3 GCE

Centre Number | Candidate Number

Chemistry

Advanced

Paper 1: Advanced Inorganic and Physical Chemistry

Tuesday 13 June 2017 – Afternoon
Time: 1 hour 45 minutes

Paper Reference
9CH0/01

You must have:
Data Booklet
Scientific calculator, ruler

Total Marks

Instructions

- Use **black** ink or **black** ball-point pen.
- **Fill in the boxes** at the top of this page with your name, centre number and candidate number.
- Answer **all** questions.
- Answer the questions in the spaces provided
  – *there may be more space than you need.*

Information

- The total mark for this paper is 90.
- The marks for **each** question are shown in brackets
  – *use this as a guide as to how much time to spend on each question.*
- You may use a scientific calculator.
- For the question marked with an **asterisk** (*), marks will be awarded for your ability to structure your answer logically showing the points that you make are related or follow on from each other where appropriate.
- A Periodic Table is printed on the back cover of this paper.

Advice

- Read each question carefully before you start to answer it.
- Try to answer every question.
- Check your answers if you have time at the end.
- Show all your working in calculations and include units where appropriate.

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Answer ALL questions.

Some questions must be answered with a cross in a box ☒.
If you change your mind about an answer, put a line through the box ☒ and then mark your new answer with a cross ☒.

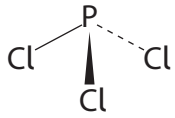
1 A phosphorus atom has mass number 31.

(a) How many of each sub-atomic particle are present in the phosphide ion, P^{3-}?

(1)

	Number of protons	Number of neutrons	Number of electrons
☒ **A**	15	16	12
☒ **B**	15	16	18
☒ **C**	16	15	12
☒ **D**	16	15	18

(b) Phosphorus(III) chloride molecules are pyramidal with a bond angle less than 109.5°.



(i) Explain why a phosphorus(III) chloride molecule has this shape and bond angle.

(2)

There are 4 pairs of electrons around the central P atom: 3 bond pairs and 1 lone pair.
These electron pairs repel and are arranged as far apart as possible to minimise repulsion, giving a pyramidal shape.
The bond angle is less than 109.5° because lone pair-bond pair repulsion is greater than bond pair-bond pair repulsion.

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(ii) Which describes the polarity of the P—Cl bond and the polarity of the phosphorus(III) chloride molecule?

(1)

	Polarity of P—Cl bond	Polarity of molecule
☐ **A**	non-polar	non-polar
☐ **B**	non-polar	polar
☐ **C**	polar	non-polar
☒ **D**	polar	polar

(c) Phosphorus has one naturally occurring isotope with mass number 31. Chlorine exists as two isotopes with mass numbers 35 and 37.

Give the formulae and mass/charge ratio of the ions responsible for the molecular ion peaks in the mass spectrum of phosphorus(III) chloride, PCl_3.

(2)

$P(^{35}Cl)_3^+$ m/z = 136

$P(^{35}Cl)_2(^{37}Cl)^+$ m/z = 138

$P(^{35}Cl)(^{37}Cl)_2^+$ m/z = 140

$P(^{37}Cl)_3^+$ m/z = 142

(Total for Question 1 = 6 marks)

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2 Magnesium nitrate decomposes on heating as shown by the equation.

$$2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$$

(a) Explain, in terms of all the relevant oxidation numbers, why this is a redox reaction.

(3)

Nitrogen: oxidation number changes from +5 in $Mg(NO_3)_2$ to +4 in NO_2, so nitrogen is reduced (oxidation number decreases).

Oxygen: oxidation number changes from −2 in $Mg(NO_3)_2$ to 0 in O_2, so oxygen is oxidised (oxidation number increases).

Mg stays +2. One element is reduced and another is oxidised, so it is a redox reaction.

(b) Calcium nitrate decomposes in a similar way to magnesium nitrate, but requires a higher temperature for decomposition.

Explain this observation in terms of the charge and size of the cations.

(3)

Both cations have the same 2+ charge, but the calcium ion, Ca^{2+}, has a larger ionic radius than the magnesium ion, Mg^{2+}.

So Ca^{2+} has a lower charge density and causes less polarisation (distortion) of the nitrate ion, NO_3^-, electron cloud.

The N–O bonds in the nitrate ion are weakened less, so more energy (a higher temperature) is needed to decompose calcium nitrate.

(Total for Question 2 = 6 marks)

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3 This question is about halogens and redox reactions.

(a) The boiling temperatures of three halogens are shown in the table.

Halogen	Boiling temperature / °C
chlorine	–35
bromine	59
iodine	184

Explain why the boiling temperatures increase from chlorine to iodine.

(2)

From chlorine to iodine the number of electrons in each molecule increases (from 34 in Cl_2 to 106 in I_2).

So the London forces (instantaneous dipole-induced dipole forces) between the molecules get stronger, and more energy is needed to separate the molecules.

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(b) Potassium halides react with concentrated sulfuric acid to form potassium hydrogensulfate and the different products shown in the table.

Potassium halide	Products
potassium chloride	hydrogen chloride
potassium bromide	hydrogen bromide, bromine and sulfur dioxide
potassium iodide	hydrogen iodide, iodine, hydrogen sulfide and sulfur

By referring to any changes in oxidation numbers when these halides react with concentrated sulfuric acid, explain which halide is the strongest reducing agent.

(3)

Iodide ions, I^-, are the strongest reducing agent.

I^- reduces the sulfur in H_2SO_4 from +6 to 0 in S and to −2 in H_2S (iodide is oxidised from −1 to 0 in I_2).

Br^- only reduces sulfur from +6 to +4 in SO_2 (bromide is oxidised from −1 to 0 in Br_2).

Cl^- does not reduce sulfuric acid at all: the oxidation numbers of Cl (−1) and S (+6) do not change, so it is not a redox reaction.

Iodide causes the largest decrease in the oxidation number of sulfur, so it is the strongest reducing agent.

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(c) Use these electrode potentials to answer the following questions.

Electrode reaction	$E^{\ominus}$ / V
$I_2(aq) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	+1.23
$Cl_2(aq) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51

(i) Which species will oxidise $Fe^{2+}(aq)$ to $Fe^{3+}(aq)$?

(1)

☒ **A** $Br_2(aq)$

☐ **B** $Cl^-(aq)$

☐ **C** $I_2(aq)$

☐ **D** $Mn^{2+}(aq)$

(ii) Write the ionic equation and calculate the $E^{\ominus}_{cell}$ value for the reaction between MnO_4^- ions and Br^- ions in acidic solution.
State symbols are not required.

(3)

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O \quad (\times 2)$

$2Br^- \rightarrow Br_2 + 2e^- \quad (\times 5)$

$2MnO_4^- + 16H^+ + 10Br^- \rightarrow 2Mn^{2+} + 8H_2O + 5Br_2$

$E^{\ominus}_{cell} = 1.51 - 1.09 = +0.42\ V$

(Total for Question 3 = 9 marks)

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4 Iron and zinc are in the d-block of the Periodic Table.

(a) Which of these is the electronic configuration of an iron(II) ion, Fe^{2+}?

(1)

		3d					4s
☒	**A** [Ar]	↑↓	↑↓	↑↓			
☒	**B** [Ar]	↑↓	↑	↑	↑	↑	
☒	**C** [Ar]	↑↓	↑↓				↑↓
☒	**D** [Ar]	↑	↑	↑	↑		↑↓

(b) Iron(II) ions, $[Fe(H_2O)_6]^{2+}$, form a pale green solution but zinc ions, $[Zn(H_2O)_6]^{2+}$, form a colourless solution.

Explain why zinc ions are colourless.

(2)

Zn^{2+} ions have a full 3d sub-shell ($3d^{10}$): all the 3d orbitals are full.

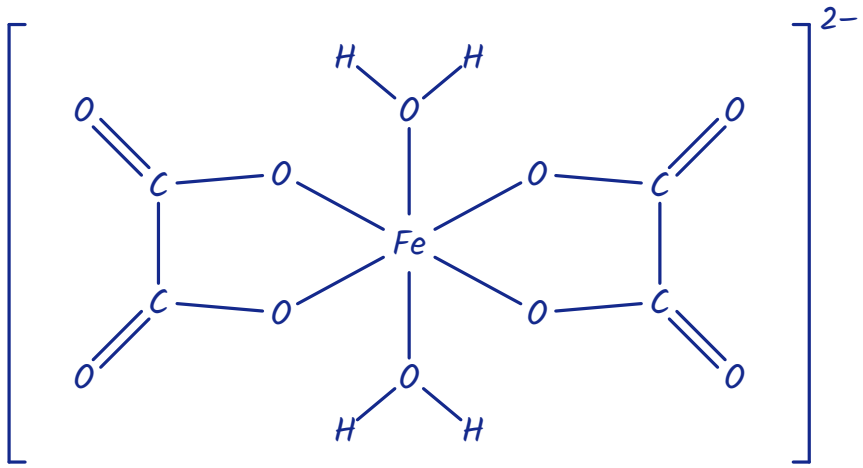
So no d-d transitions can take place: electrons cannot be promoted to a higher energy 3d orbital, so no visible light is absorbed.

(c) Hydrated iron(II) ions react with ethanedioate ions, $C_2O_4^{2-}$, to form a complex ion.

$$[Fe(H_2O)_6]^{2+} + 2C_2O_4^{2-} \rightleftharpoons [Fe(C_2O_4)_2(H_2O)_2]^{2-} + 4H_2O$$

(i) Draw a structure of the $[Fe(C_2O_4)_2(H_2O)_2]^{2-}$ ion, showing **all** of the bonds.

(2)



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(ii) Explain, in terms of entropy, why this reaction is feasible.

(2)

There are 3 particles (moles) on the left of the equation and 5 on the right, so the number of particles increases.

So the entropy of the system increases (ΔS_{system} is positive) and $\Delta S_{surroundings}$ is about unchanged, so ΔS_{total} is positive and the reaction is feasible.

(d) Iodide ions, I^-, react with peroxodisulfate(VI) ions, $S_2O_8^{2-}$

$$2I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_2(aq) + 2SO_4^{2-}(aq)$$

This reaction is catalysed by iron(II) ions, $Fe^{2+}(aq)$.

Write **two** ionic equations to show how iron(II) ions act as a catalyst in this reaction. State symbols are not required.

(2)

$2Fe^{2+} + S_2O_8^{2-} \rightarrow 2Fe^{3+} + 2SO_4^{2-}$

$2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$

(Total for Question 4 = 9 marks)

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5 This question is about enthalpy changes and entropy changes.

(a) Which is the equation for the standard enthalpy change of formation, $\Delta_f H^{\ominus}$, of aluminium oxide?

(1)

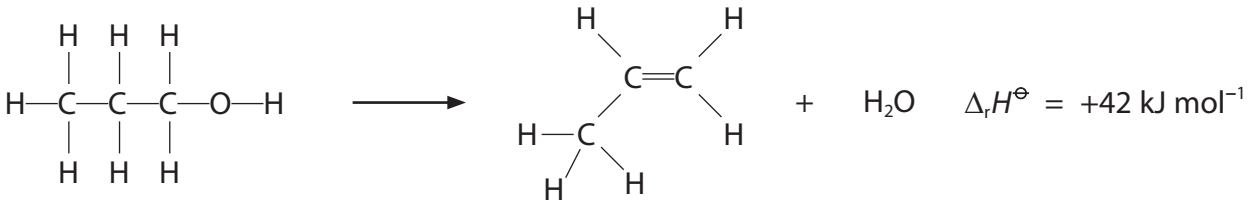
☒ **A** $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

☒ **B** $4Al(s) + 6O(g) \rightarrow 2Al_2O_3(s)$

☒ **C** $2Al(s) + 1½O_2(g) \rightarrow Al_2O_3(s)$

☒ **D** $2Al(s) + 3O(g) \rightarrow Al_2O_3(s)$

(b) Propan-1-ol is dehydrated to form propene.



The relevant mean bond enthalpies are given in the table.

Bond	Mean bond enthalpy / kJ mol^{-1}
C—C	347
C═C	612
C—H	413
O—H	464

Calculate the C—O mean bond enthalpy, using the mean bond enthalpies given in the table and the enthalpy change of reaction.

(3)

Bonds broken: C–C + C–H + C–O = 347 + 413 + (C–O) = (C–O) + 760

Bonds made: C=C + O–H = 612 + 464 = 1076

ΔH = bonds broken − bonds made

(C–O) + 760 − 1076 = +42

C–O = 42 + 1076 − 760 = +358 kJ mol^{-1}

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(c) Which reaction has a negative value for ΔS_{system}?

(1)

☒ **A** $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

☐ **B** $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$

☐ **C** $MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$

☐ **D** $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

(d) What is the expression for ΔS_{total}?

(1)

☐ **A** $\Delta S_{surroundings} + \frac{\Delta H}{T}$

☐ **B** $\Delta S_{surroundings} - \frac{\Delta H}{T}$

☐ **C** $\Delta S_{system} + \frac{\Delta H}{T}$

☒ **D** $\Delta S_{system} - \frac{\Delta H}{T}$

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(e) Calcium carbonate decomposes on heating.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

$\Delta_r H = +178 \text{ kJ mol}^{-1}$

$\Delta S_{system} = +165 \text{ J mol}^{-1} \text{ K}^{-1}$

Show, by calculating the value for the free energy change, ΔG, that this decomposition is not feasible at 298 K, and then calculate the minimum temperature to which calcium carbonate must be heated to make it decompose.

(3)

$\Delta G = \Delta H - T\Delta S_{system}$

$\Delta G = 178 - (298 \times 0.165)$

$\Delta G = +128.8 \text{ kJ mol}^{-1}$ $(+129 \text{ kJ mol}^{-1})$

ΔG is positive (> 0), so the decomposition is not feasible at 298 K.

Minimum temperature is when $\Delta G = 0$, so $T = \Delta H \div \Delta S_{system}$

$T = 178 \div 0.165 = 1079 \text{ K } (806 \text{ °C})$

(Total for Question 5 = 9 marks)

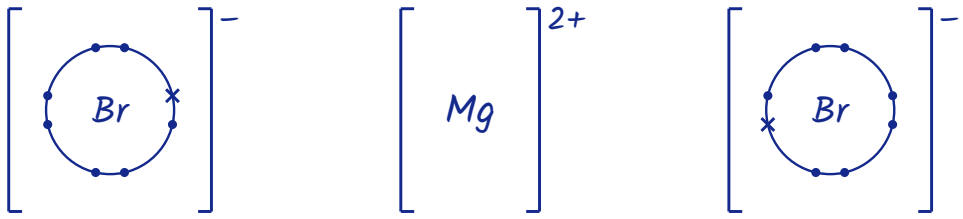
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6 Magnesium bromide, $MgBr_2$, is an ionic compound.

(a) (i) Draw a dot-and-cross diagram to show the bonding in magnesium bromide.
Only outer shell electrons are required.

(1)



(· = Br electron, × = electron from Mg)

(ii) State all the conditions under which magnesium bromide conducts electricity.

(1)

When molten (liquid) and when dissolved in water (aqueous solution).

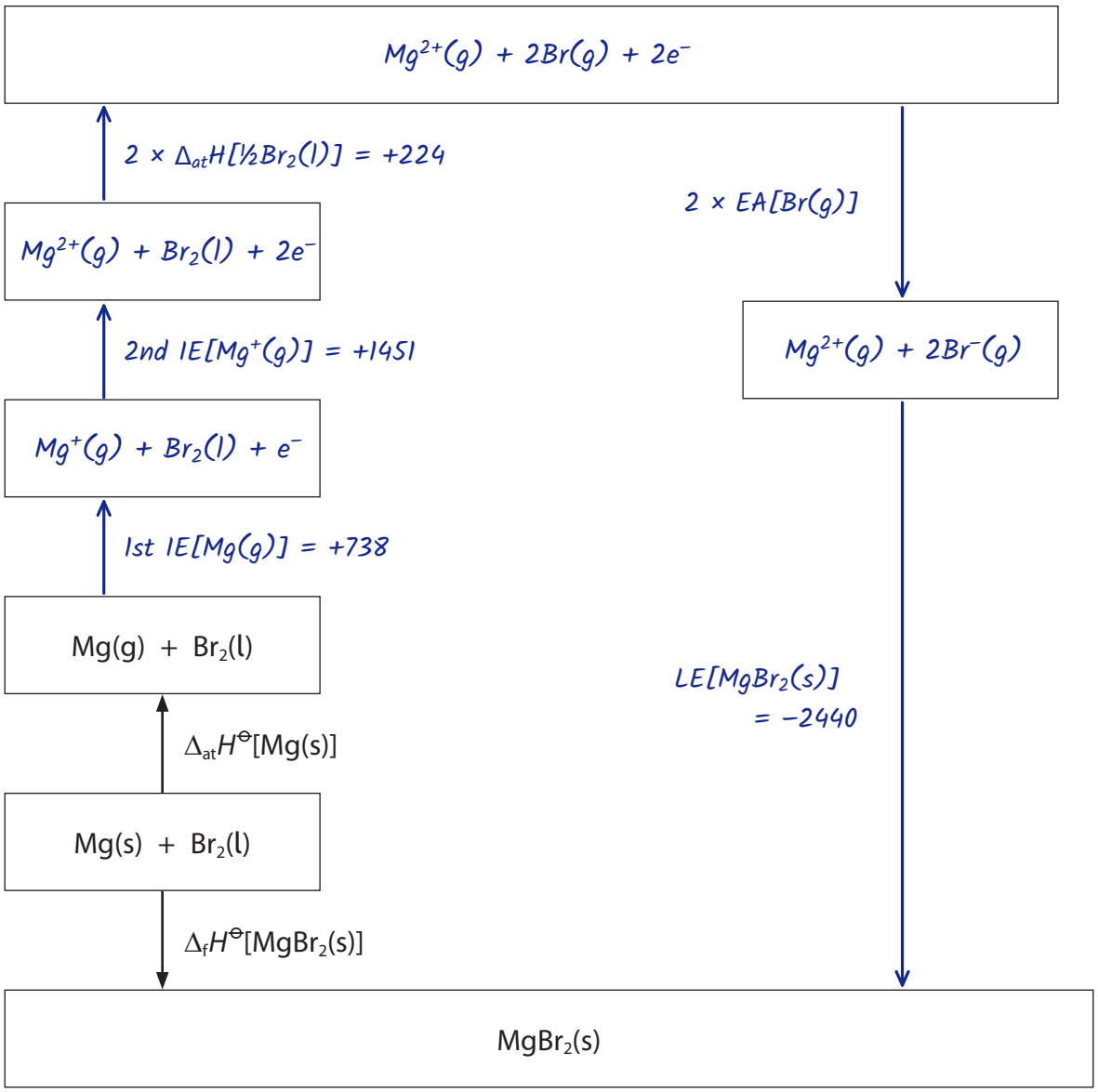
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(b) The table shows the enthalpy changes needed to calculate the first electron affinity of bromine.

Enthalpy change	Value / kJ mol^{-1}
enthalpy change of atomisation of magnesium, $\Delta_{at}H^{\ominus}$[Mg(s)]	+148
1st ionisation energy of magnesium, 1st IE[Mg(g)]	+738
2nd ionisation energy of magnesium, 2nd IE[Mg^{+}(g)]	+1 451
enthalpy change of atomisation of bromine, $\Delta_{at}H^{\ominus}$[½Br_2(l)]	+112
lattice energy of magnesium bromide, LE[$MgBr_2$(s)]	–2 440
enthalpy change of formation of magnesium bromide, $\Delta_f H^{\ominus}$[$MgBr_2$(s)]	–524

(i) Complete the Born-Haber cycle for magnesium bromide with formulae, electrons and labelled arrows. The cycle is not drawn to scale.

(3)



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(ii) Calculate the first electron affinity of bromine, in kJ mol^{-1}.

(2)

$2 \times EA[Br] = \Delta_f H - \Delta_{at}H[Mg] - 1st\ IE - 2nd\ IE - 2\Delta_{at}H[\frac{1}{2}Br_2] - LE$

$2 \times EA = -524 - (+148) - (+738) - (+1451) - (+224) - (-2440)$

$2 \times EA = -645$

$EA[Br(g)] = -645 \div 2 = -322.5\ kJ\ mol^{-1}$

(c) (i) The first ionisation energy of sodium is 496 kJ mol^{-1}.

Explain why the first ionisation energy of magnesium is higher than that of sodium.

(3)

Magnesium has more protons than sodium (12 compared with 11), so it has a greater nuclear charge.

The outer electron in Mg is in the same shell (3s sub-shell) as the outer electron in Na, so the shielding by inner electrons is similar.

So the force of attraction between the nucleus and the outer electron is greater in magnesium than in sodium, so more energy is needed to remove it.

(ii) Write the equation, including state symbols, to show the **third** ionisation energy of magnesium.

(1)

$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^-$

(Total for Question 6 = 11 marks)

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7 In acid-base neutralisation reactions, there is a temperature change.

(a) The enthalpy change when hydrochloric acid reacts with aqueous ammonia is –53.4 kJ mol^{-1}.

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$

Calculate the temperature change you would expect when 25.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid is mixed with 25.0 cm^3 of 1.00 mol dm^{-3} aqueous ammonia.

Give your answer to an appropriate number of significant figures.

[Assume: the density of the solution is 1.00 g cm^{-3}
the specific heat capacity of the solution is 4.18 J g^{-1} $°C^{-1}$]

(3)

Moles of HCl = 25.0 × 1.00 ÷ 1000 = 0.0250 mol

Energy released = 0.0250 × 53.4 = 1.335 kJ = 1335 J

Q = mcΔT, mass of solution = 50.0 g

ΔT = 1335 ÷ (50.0 × 4.18) = 6.39

Temperature change = 6.4 °C (a rise), to 2 s.f.

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*(b) The table shows the enthalpy changes of reaction when 1 mol of different acids are neutralised by sodium hydroxide solution, at 298 K.

Acid	Enthalpy change of reaction for 1 mol of acid / kJ mol^{-1}
hydrochloric acid, HCl	–58
nitric acid, HNO_3	–58
sulfuric acid, H_2SO_4	–115
ethanoic acid, CH_3COOH	–56

Comment on the relative enthalpy changes of reaction, using the data from the table and including any relevant equations.

(6)

Hydrochloric acid and nitric acid have the same value, −58 kJ mol^{-1}, because both are strong acids, completely dissociated into ions in solution.

Both are monoprotic, so 1 mol of acid provides 1 mol of H^+ and forms 1 mol of water.

The reaction in each case is the same: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.

Sulfuric acid is diprotic (dibasic): $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$.

So 1 mol of sulfuric acid provides 2 mol of H^+ and forms 2 mol of water, so its value, −115 kJ mol^{-1}, is almost twice that of hydrochloric or nitric acid.

Ethanoic acid is a weak acid, only partially dissociated: $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$.

So some energy is needed to break the O–H bond to ionise the ethanoic acid and release H^+ ions, so its enthalpy change, −56 kJ mol^{-1}, is less exothermic than for the strong acids.

(Total for Question 7 = 9 marks)

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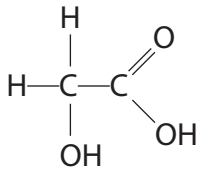
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8 2-Hydroxyethanoic acid, also known as glycolic acid, $CH_2OHCOOH$, is an alpha hydroxy acid used in some skincare products.
It has a K_a value of 1.5×10^{-4} mol dm^{-3}.

The structure of glycolic acid is



(a) A solution of glycolic acid of concentration 0.1 mol dm^{-3} has a pH of 2.4

What is the approximate pH of the resulting solution after it has been diluted by a factor of 100?

(1)

☒ **A** 1.4

☒ **B** 2.4

☒ **C** 3.4

☒ **D** 4.4

(b) Another solution of glycolic acid has a pH of 2.0

Calculate the concentration of this solution.

(3)

$[H^+] = 10^{-pH} = 10^{-2.0} = 0.0100$ mol dm^{-3}

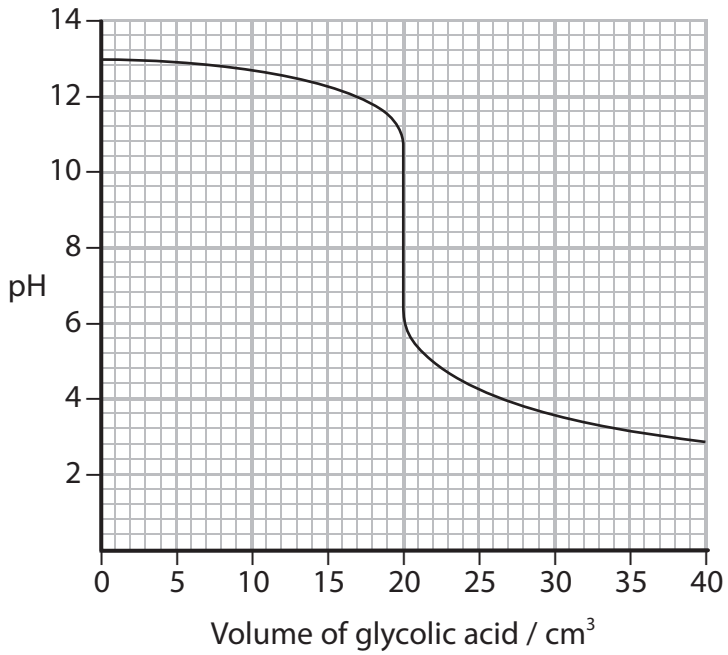
$K_a = [H^+]^2 \div [CH_2OHCOOH]$

$[CH_2OHCOOH] = [H^+]^2 \div K_a = (0.0100)^2 \div (1.5 \times 10^{-4})$

$[CH_2OHCOOH] = 0.67$ mol dm^{-3}

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(c) The titration curve for adding glycolic acid to 25.0 cm^3 of 0.100 mol dm^{-3} sodium hydroxide is shown.



(i) Use the information given in your Data Booklet to select a suitable indicator for this titration, including the colour change you would expect to see.

Justify your selection.

(3)

Phenolphthalein is a suitable indicator.

The colour change would be from pink to colourless (as the acid is added to the alkali).

Its pH range (8.2 to 10.0) lies completely within the vertical section of the curve (about pH 6 to 11), as this is a titration of a weak acid with a strong base.

(ii) What is the concentration of this glycolic acid in mol dm^{-3}?

(1)

☐ **A** 0.080

☐ **B** 0.100

☒ **C** 0.125

☐ **D** 0.250

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(iii) The pH of the solution containing just sodium glycolate and water is

(1)

- ☐ **A** 2.8
- ☐ **B** 6.0
- ☒ **C** 8.3
- ☐ **D** 11.0

(d) Glycolic acid has an acid dissociation constant of 1.5×10^{-4} mol dm^{-3} compared with a value of 1.7×10^{-5} mol dm^{-3} for ethanoic acid.

(i) Give a possible explanation as to why the value of K_a for glycolic acid is approximately ten times larger than that of ethanoic acid.

(2)

The oxygen atom of the extra OH (hydroxyl) group on the alpha carbon is electronegative, so it withdraws electrons from the COOH group. This weakens the O–H bond in the COOH group and stabilises the CH_2OHCOO^- anion, so H^+ is lost more easily.

(ii) Complete the equation to show the conjugate acid-base pairs that would be produced when pure samples of glycolic acid and ethanoic acid are mixed.

(1)

$CH_2OHCOOH + CH_3COOH \rightarrow$ *CH_2OHCOO^-* + *$CH_3COOH_2^+$*

(Total for Question 8 = 12 marks)

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9 This question is about buffer solutions.

(a) A buffer solution is formed from disodium hydrogenphosphate, containing HPO_4^{2-} ions, and sodium dihydrogenphosphate, containing $H_2PO_4^-$ ions.

Write the **ionic** equations involving HPO_4^{2-} and $H_2PO_4^-$ ions to show how this solution acts as a buffer solution.

(2)

$HPO_4^{2-} + H^+ \rightarrow H_2PO_4^-$ (removes added acid)

$H_2PO_4^- + OH^- \rightarrow HPO_4^{2-} + H_2O$ (removes added alkali)

(b) Another buffer solution was formed by mixing 20.0 cm^3 of sodium hydroxide solution of concentration 0.100 mol dm^{-3} with 25.0 cm^3 of ethanoic acid of concentration 0.150 mol dm^{-3}.

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

Calculate the pH of this buffer solution.

[K_a for ethanoic acid = 1.74×10^{-5} mol dm^{-3}]

(5)

$n(NaOH) = n(CH_3COONa) = 0.100 \times 20.0 \div 1000 = 0.00200$ mol

initial $n(CH_3COOH) = 0.150 \times 25.0 \div 1000 = 0.00375$ mol

acid left = 0.00375 − 0.00200 = 0.00175 mol

total volume = 45.0 cm^3, so [salt] = 0.0444 and [acid] = 0.0389 mol dm^{-3}

$[H^+] = K_a[acid] \div [salt] = 1.74 \times 10^{-5} \times 0.0389 \div 0.0444$

$[H^+] = 1.52 \times 10^{-5}$ mol dm^{-3}

$pH = -\log(1.52 \times 10^{-5}) = 4.82$

(Total for Question 9 = 7 marks)

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10 Hydrogen is produced on a large scale by several different processes.

(a) One process for producing hydrogen involves reacting white-hot carbon with steam.

$$C(s) + H_2O(g) \rightleftharpoons H_2(g) + CO(g) \qquad \Delta H = +131 \text{ kJ mol}^{-1}$$

The expression for the equilibrium constant, K_p, is

$$K_p = \frac{p(H_2)\, p(CO)}{p(H_2O)}$$

(i) Give a reason why the partial pressure of carbon is not included in the expression.

(1)

Carbon is a solid, so it has no (constant) partial pressure and does not contribute to the pressure of the system.

(ii) Explain the effect of an increase in pressure on the equilibrium position of this reaction.

(2)

There are fewer moles of gas on the left (1 mol of H_2O) than on the right (2 mol: H_2 and CO).

So an increase in pressure shifts the equilibrium position to the left, the side with fewer gas molecules.

(iii) Explain, by reference to any change in the value of K_p, the effect of an increase in temperature on the equilibrium position of this reaction.

(2)

The forward reaction is endothermic ($\Delta H = +131$ kJ mol^{-1}), so K_p increases as the temperature increases.

So the equilibrium position shifts to the right, towards the products H_2 and CO.

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(iv) At 1000 K and a total pressure of 2.0 atm, 1.00 mol of steam reacted with excess carbon.

At equilibrium, 0.81 mol of hydrogen was present.

Calculate the value of K_p at 1000 K, stating any units.

(4)

At equilibrium: $n(H_2) = n(CO) = 0.81$ mol

$n(H_2O) = 1.00 - 0.81 = 0.19$ mol

Total moles $= 0.19 + 0.81 + 0.81 = 1.81$ mol

$p(H_2O) = (0.19 \div 1.81) \times 2.0 = 0.210$ atm

$p(H_2) = p(CO) = (0.81 \div 1.81) \times 2.0 = 0.895$ atm

$K_p = p(H_2)\, p(CO) \div p(H_2O) = (0.895 \times 0.895) \div 0.210$

$K_p = 3.82$ atm

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(b) Carbon monoxide reacts with steam.

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

At 1100 K, $K_c = 1.00$

In an experiment, 1 mol of carbon monoxide was mixed with 1 mol of steam, 2 mol of carbon dioxide and 2 mol of hydrogen.

Deduce, with reasons, the direction in which the reaction will shift to reach equilibrium.

(3)

$Q = [CO_2][H_2] \div [CO][H_2O] = (2 \times 2) \div (1 \times 1) = 4$

$Q = 4$ is larger than $K_c = 1.00$, so the concentrations of CO_2 and H_2 (products) must decrease and those of CO and H_2O (reactants) must increase, until there is 1.5 mol of each.

So the reaction shifts to the left (the reverse reaction is favoured).

(Total for Question 10 = 12 marks)

TOTAL FOR PAPER = 90 MARKS

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The Periodic Table of Elements

1	2											3	4	5	6	7	0 (8)
							1.0 H hydrogen 1										(18) 4.0 He helium 2
(1)	(2)		Key									(13)	(14)	(15)	(16)	(17)	
6.9 Li lithium 3	9.0 Be beryllium 4		relative atomic mass atomic symbol name atomic (proton) number									10.8 B boron 5	12.0 C carbon 6	14.0 N nitrogen 7	16.0 O oxygen 8	19.0 F fluorine 9	20.2 Ne neon 10
23.0 Na sodium 11	24.3 Mg magnesium 12	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	27.0 Al aluminium 13	28.1 Si silicon 14	31.0 P phosphorus 15	32.1 S sulfur 16	35.5 Cl chlorine 17	39.9 Ar argon 18
39.1 K potassium 19	40.1 Ca calcium 20	45.0 Sc scandium 21	47.9 Ti titanium 22	50.9 V vanadium 23	52.0 Cr chromium 24	54.9 Mn manganese 25	55.8 Fe iron 26	58.9 Co cobalt 27	58.7 Ni nickel 28	63.5 Cu copper 29	65.4 Zn zinc 30	69.7 Ga gallium 31	72.6 Ge germanium 32	74.9 As arsenic 33	79.0 Se selenium 34	79.9 Br bromine 35	83.8 Kr krypton 36
85.5 Rb rubidium 37	87.6 Sr strontium 38	88.9 Y yttrium 39	91.2 Zr zirconium 40	92.9 Nb niobium 41	95.9 Mo molybdenum 42	[98] Tc technetium 43	101.1 Ru ruthenium 44	102.9 Rh rhodium 45	106.4 Pd palladium 46	107.9 Ag silver 47	112.4 Cd cadmium 48	114.8 In indium 49	118.7 Sn tin 50	121.8 Sb antimony 51	127.6 Te tellurium 52	126.9 I iodine 53	131.3 Xe xenon 54
132.9 Cs caesium 55	137.3 Ba barium 56	138.9 La* lanthanum 57	178.5 Hf hafnium 72	180.9 Ta tantalum 73	183.8 W tungsten 74	186.2 Re rhenium 75	190.2 Os osmium 76	192.2 Ir iridium 77	195.1 Pt platinum 78	197.0 Au gold 79	200.6 Hg mercury 80	204.4 Tl thallium 81	207.2 Pb lead 82	209.0 Bi bismuth 83	[209] Po polonium 84	[210] At astatine 85	[222] Rn radon 86
[223] Fr francium 87	[226] Ra radium 88	[227] Ac* actinium 89	[261] Rf rutherfordium 104	[262] Db dubnium 105	[266] Sg seaborgium 106	[264] Bh bohrium 107	[277] Hs hassium 108	[268] Mt meitnerium 109	[271] Ds damstadtium 110	[272] Rg roentgenium 111							

Elements with atomic numbers 112-116 have been reported but not fully authenticated

* Lanthanide series	140 Ce cerium 58	141 Pr praseodymium 59	144 Nd neodymium 60	[147] Pm promethium 61	150 Sm samarium 62	152 Eu europium 63	157 Gd gadolinium 64	159 Tb terbium 65	163 Dy dysprosium 66	165 Ho holmium 67	167 Er erbium 68	169 Tm thulium 69	173 Yb ytterbium 70	175 Lu lutetium 71
* Actinide series	232 Th thorium 90	[231] Pa protactinium 91	238 U uranium 92	[237] Np neptunium 93	[242] Pu plutonium 94	[243] Am americium 95	[247] Cm curium 96	[245] Bk berkelium 97	[251] Cf californium 98	[254] Es einsteinium 99	[253] Fm fermium 100	[256] Md mendelevium 101	[254] No nobelium 102	[257] Lr lawrencium 103

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