Please check the examination details below before entering your candidate information

Candidate surname	Other names
StemUp Model Solution	

Pearson Edexcel Level 3 GCE

Centre Number	Candidate Number

Tuesday 4 June 2019

Afternoon (Time: 1 hour 45 minutes) | Paper Reference **9CH0/01**

Chemistry

Advanced

Paper 1: Advanced Inorganic and Physical Chemistry

Candidates 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.*
- 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.
- Show all your working in calculations and include units where appropriate.
- Check your answers if you have time at the end.

Turn over ▶

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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 This question is about atoms, molecules and ions.

(a) The numbers of subatomic particles in an ^{18}O atom are

(1)

☒ **A** 8 protons, 10 neutrons and 8 electrons

☐ **B** 9 protons, 9 neutrons and 9 electrons

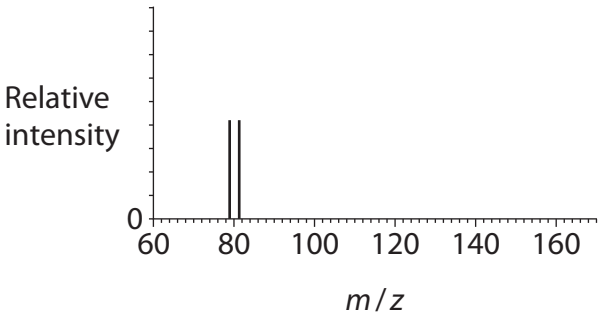
☐ **C** 10 protons, 8 neutrons and 10 electrons

☐ **D** 18 protons, 18 neutrons and 18 electrons

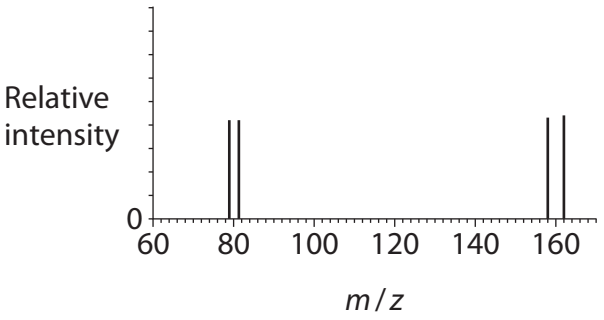
(b) The mass spectrum of a sample of bromine **molecules** with approximately equal proportions of the ^{79}Br and ^{81}Br isotopes is

(1)

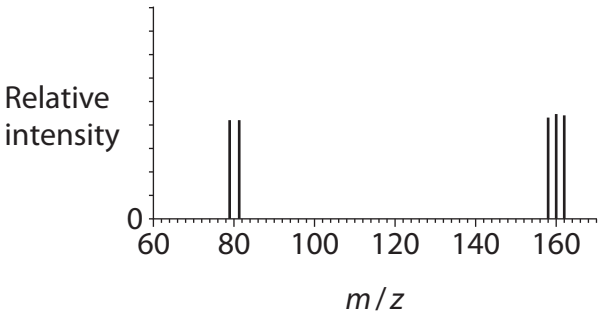
☐ **A**



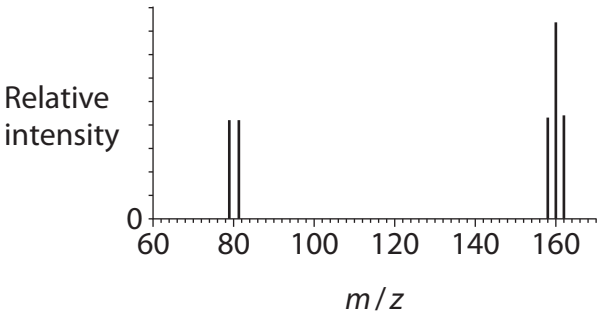
☐ **B**



☐ **C**



☒ **D**



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(c) The total number of electrons in **all** the occupied **p** orbitals in a chloride ion, Cl^-, is

(1)

☒ **A** 5

☒ **B** 6

☒ **C** 12

☒ **D** 18

(d) Which of these isoelectronic ions has the largest ionic radius?

(1)

☒ **A** N^{3-}

☒ **B** O^{2-}

☒ **C** Na^+

☒ **D** Al^{3+}

(Total for Question 1 = 4 marks)

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2 This question is about some redox reactions of chlorine, bromine and iodine.

(a) An **excess** of aqueous potassium bromide was added to chlorine water and the solution turned orange.

(i) Write an equation for this reaction. State symbols are not required.

(1)

$Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$

(ii) Silver nitrate solution was added to the mixture in (a) and excess dilute ammonia solution was then added to the precipitate formed. Only some of the precipitate dissolved.

Deduce why only **some** of the precipitate dissolved.

(3)

The precipitate is a mixture of silver chloride and silver bromide, because the bromide ions were in excess, so not all of the bromide ions were oxidised by the chlorine.

Silver chloride, AgCl, dissolves in dilute ammonia solution.

Silver bromide, AgBr, does not dissolve in dilute ammonia (it only dissolves in concentrated ammonia), so only some of the precipitate dissolves.

(iii) Aqueous potassium bromide was added to aqueous iodine, instead of chlorine water. There was no reaction.

Give a reason why no reaction occurred.

(1)

Iodine is a weaker oxidising agent than chlorine (and bromine), so iodine cannot oxidise bromide ions.

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(b) Chlorine undergoes disproportionation when it reacts with **hot** aqueous sodium hydroxide solution.

(i) Complete the ionic equation for this reaction.
State symbols are not required.

(1)

3 Cl_2 + 6 OH^- → 5 Cl^- + 1 ClO_3^- + 3 H_2O

(ii) Explain, in terms of oxidation numbers, why this is a disproportionation reaction.

(2)

Chlorine is both reduced and oxidised in the same reaction.

The oxidation number of chlorine changes from 0 in Cl_2 to −1 in Cl^-, so chlorine is reduced.

The oxidation number of chlorine also changes from 0 in Cl_2 to +5 in ClO_3^-, so chlorine is oxidised.

(Total for Question 2 = 8 marks)

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3 The standard molar enthalpy change of neutralisation is the enthalpy change when an acid and an alkali react under standard conditions to form one mole of water.

(a) An experiment was carried out to determine the enthalpy change of neutralisation for the reaction between propanoic acid and sodium hydroxide.

The equation for this reaction is

$$CH_3CH_2COOH(aq) + NaOH(aq) \rightarrow CH_3CH_2COO^-Na^+(aq) + H_2O(l)$$

50.0 cm^3 of sodium hydroxide solution, of concentration 1.00 mol dm^{-3}, was placed in a polystyrene cup. The initial temperature was measured.

(i) Which piece of equipment has the **smallest** measurement uncertainty for the measurement of 50.0 cm^3 of sodium hydroxide solution?

(1)

		Equipment	Measurement uncertainty for **each** reading
☒	**A**	burette	±0.05 cm^3
☒	**B**	50 cm^3 measuring cylinder	±1 cm^3
☒	**C**	25 cm^3 pipette	±0.06 cm^3
☒	**D**	50 cm^3 pipette	±0.08 cm^3

(ii) 50.0 cm^3 of propanoic acid solution, of concentration 1.00 mol dm^{-3}, was added and thoroughly mixed with the sodium hydroxide solution in the polystyrene cup.

The maximum temperature rise was 6.5 °C.

Calculate the enthalpy change of neutralisation for propanoic acid, in kJ mol^{-1}, giving your answer to the **nearest whole number**.

[Assume density of the mixture = 1.00 g cm^{-3}, specific heat capacity of the mixture = 4.18 J g^{-1} $°C^{-1}$]

(3)

$Q = mc\Delta T = 100 \times 4.18 \times 6.5 = 2717\ J = 2.717\ kJ$

moles of water formed = moles of NaOH $= 0.0500 \times 1.00 = 0.0500$ mol

$\Delta H = -2.717 \div 0.0500 = -54.34\ kJ\ mol^{-1}$

$\Delta H_{neut} = -54\ kJ\ mol^{-1}$

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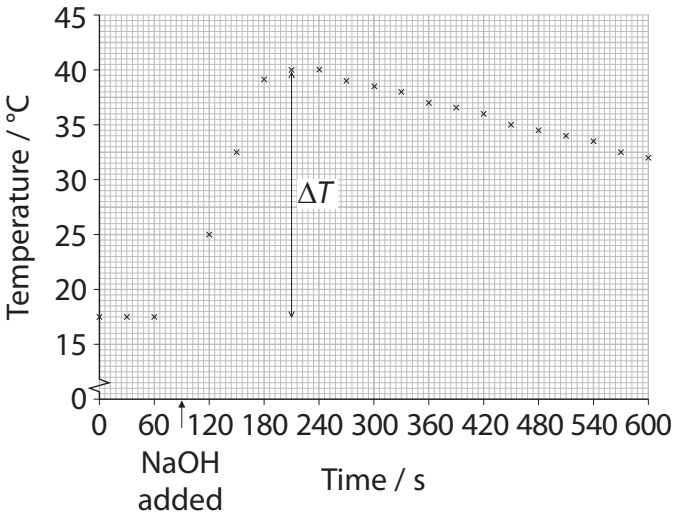
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(b) Another experiment was carried out with a solution of ethanoic acid and sodium hydroxide solution of the same concentration.

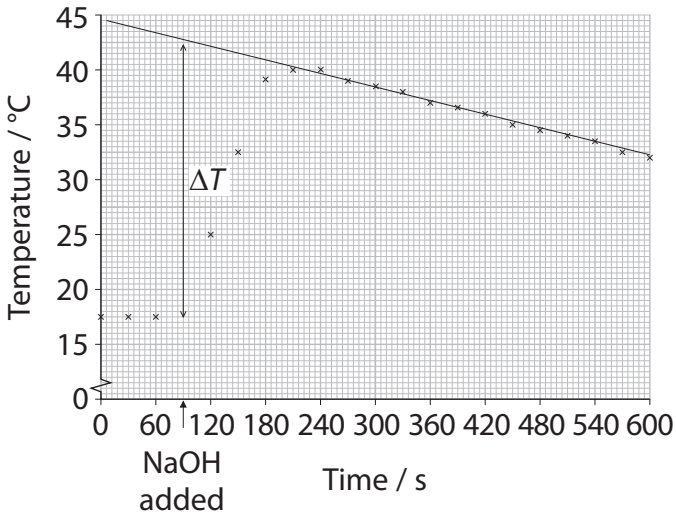
(i) Which graph shows the correct way that the maximum temperature rise should be determined?

(1)

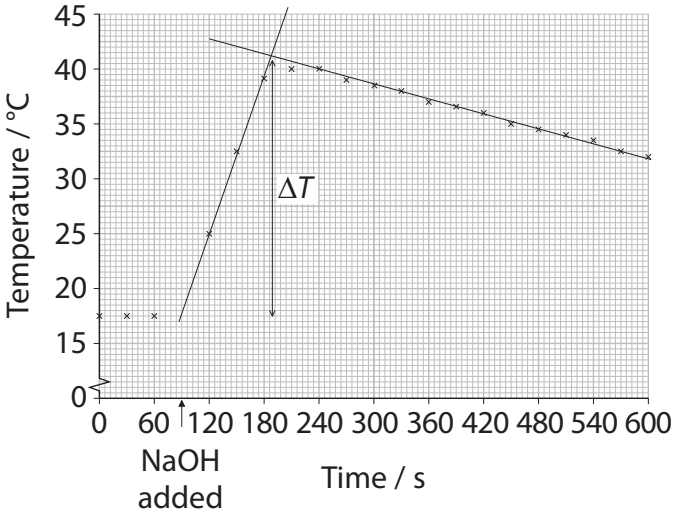
☒ **A**



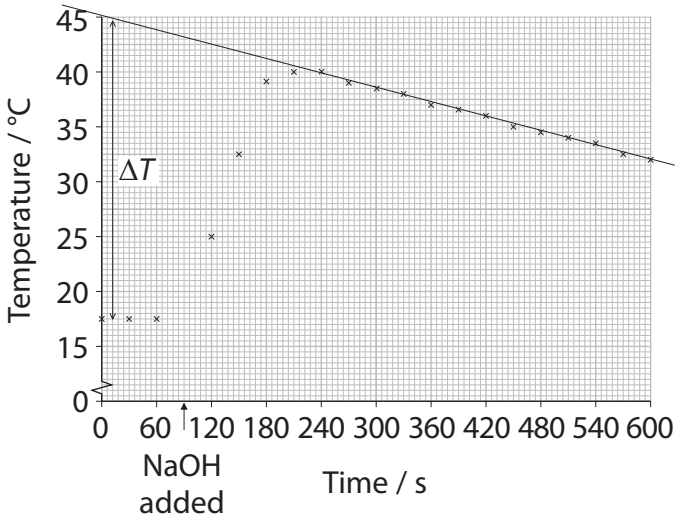
☒ **B**



☒ **C**



☒ **D**



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(ii) Explain why the data book value for the standard enthalpy change of neutralisation of ethanoic acid with sodium hydroxide is $-55.2\ kJ\ mol^{-1}$ but the value for hydrochloric acid is $-57.1\ kJ\ mol^{-1}$.

(2)

Ethanoic acid is a weak acid, so it is only partially ionised (dissociated) in solution, whereas hydrochloric acid is a strong acid and is fully ionised. Some of the energy released in the neutralisation is used to fully ionise the ethanoic acid, so less energy is released overall and the value is less exothermic.

(Total for Question 3 = 7 marks)

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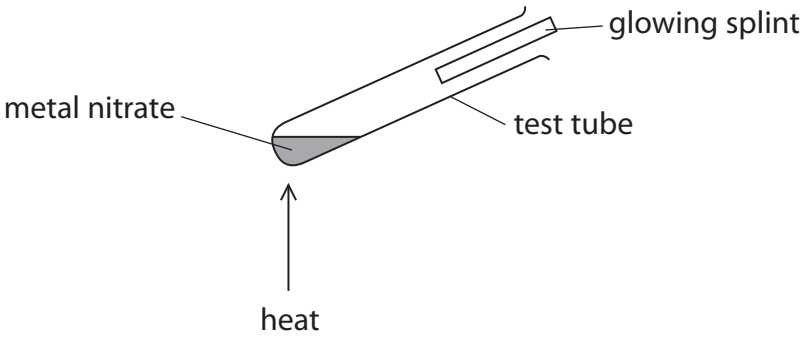
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4 Thermal decomposition is the breaking down of a substance by heat.

(a) An experiment was carried out to investigate the thermal decomposition of a metal nitrate using the apparatus shown.



(i) The glowing splint is used as a test for one of the gases given off in this experiment.

Identify this gas and the positive result of the test.

(1)

Oxygen, O_2: the glowing splint relights.

(ii) Give the name and appearance of the other gas given off in this experiment when a Group 2 nitrate is heated.

(1)

Nitrogen dioxide, NO_2: a brown gas (brown fumes).

(iii) Write the equation for the decomposition if the Group 1 compound, sodium nitrate, was used in this experiment.
State symbols are not required.

(1)

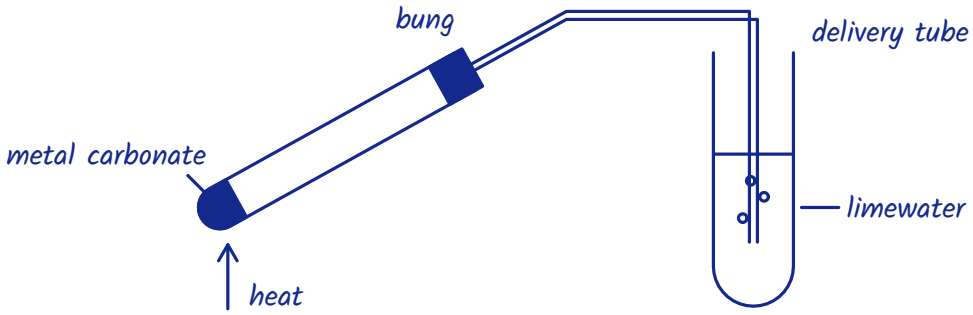
$2NaNO_3 \rightarrow 2NaNO_2 + O_2$

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(iv) Describe the apparatus that would be used to compare the decomposition of metal carbonates. Include how the rate of decomposition would be compared.

(2)



Heat a known mass of the metal carbonate in a boiling tube fitted with a bung and a delivery tube.

The delivery tube dips into limewater in a second test tube, so the carbon dioxide given off bubbles through the limewater.

Use the same mass (or moles) of each carbonate and the same flame, and time how long it takes for the limewater to turn cloudy (milky): the shorter the time, the faster the carbonate decomposes.

(b) Explain why magnesium carbonate decomposes much more readily on heating than barium carbonate.

(3)

The magnesium ion, Mg^{2+}, is smaller than the barium ion, Ba^{2+}, but has the same charge, so Mg^{2+} has a higher charge density.

The Mg^{2+} ion therefore polarises (distorts the electron cloud of) the large carbonate ion more than the Ba^{2+} ion does.

This weakens the C–O bonds in the carbonate ion more, so less energy is needed to break them and $MgCO_3$ decomposes into MgO and CO_2 at a lower temperature.

(Total for Question 4 = 8 marks)

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5 This is a question about catalysis.

(a) The reaction between iodide ions and peroxodisulfate ions is catalysed by iron(II) ions.

$$2I^- + S_2O_8^{2-} \xrightarrow{Fe^{2+}} I_2 + 2SO_4^{2-}$$

(i) Give a reason why the reaction between iodide ions and peroxodisulfate ions has a high activation energy and is therefore very slow without a catalyst.

(1)

Both reactants are negative ions (I^- and $S_2O_8^{2-}$), so they repel each other.

(ii) Explain, with the aid of two equations, how the iron(II) ions catalyse this reaction. State symbols are not required.

(3)

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

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

Fe^{2+} is first oxidised to Fe^{3+}, then Fe^{3+} is reduced back to Fe^{2+}.

Iron has variable oxidation states (+2 and +3), and each step involves oppositely charged ions that attract, so the activation energy is lower.

(b) The oxidation of sulfate(IV) ions to sulfate(VI) ions is catalysed by cobalt(II) ions in **acidic** solution. The role of cobalt(II) ions is similar to that of iron(II) ions in (a).

$$SO_3^{2-} + \tfrac{1}{2}O_2 \xrightarrow{Co^{2+}} SO_4^{2-}$$

Deduce two ionic equations to show how cobalt(II) ions catalyse the reaction in **acidic** solution. State symbols are not required.

(2)

$2Co^{2+} + \tfrac{1}{2}O_2 + 2H^+ \rightarrow 2Co^{3+} + H_2O$

$2Co^{3+} + SO_3^{2-} + H_2O \rightarrow 2Co^{2+} + SO_4^{2-} + 2H^+$

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(c) The rate of oxidation of ethanedioate ions by manganate(VII) ions starts slowly and then rapidly increases.

$$2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$$

What is the catalyst in this reaction?

(1)

☒ **A** CO_2

☒ **B** H^+

☒ **C** Mn^{2+}

☒ **D** MnO_4^-

(d) The trend in the strength of gaseous adsorption by three transition elements is

tungsten > platinum > silver

Silver is not suitable as a replacement for platinum in a catalytic converter because the adsorption of gases is too weak to allow significant chemical reaction.

Give a possible reason why tungsten would also **not** be a suitable replacement for platinum in a catalytic converter. Refer to the mechanism of heterogenous catalysis in your answer.

(1)

Tungsten adsorbs the gases too strongly, so the products would not desorb (desorption would be too slow). The active sites on the surface would stay blocked, so the reaction could not continue.

(Total for Question 5 = 8 marks)

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6 This is a question about water.

(a) Water might be expected to have a lower boiling temperature than hydrogen sulfide but it actually has a higher boiling temperature.

Comment on this statement by referring to the intermolecular forces in both these substances.

A detailed description of how the intermolecular forces arise is not required.

(4)

A lower boiling temperature might be expected for water because a water molecule has fewer electrons (10) than a hydrogen sulfide molecule (18).

So water has weaker London forces between its molecules than hydrogen sulfide.

However, water actually has a higher boiling temperature because there is hydrogen bonding between water molecules, whereas hydrogen sulfide does not form hydrogen bonds.

Hydrogen bonds are much stronger than London forces, so more energy is needed to break the hydrogen bonds between water molecules, giving water the higher boiling temperature.

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(b) Explain why both water and carbon dioxide molecules have polar bonds but only water is a polar molecule.

(4)

Oxygen is more electronegative than both hydrogen and carbon, so it has a greater attraction for the bonding pair of electrons.

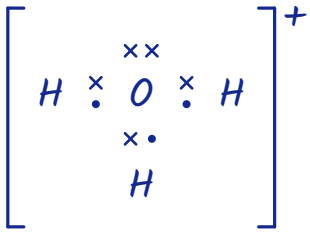
So the O–H bonds and the C=O bonds are polar, with oxygen δ– and hydrogen or carbon δ+.

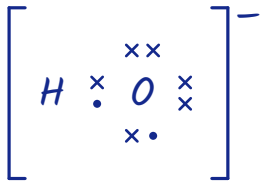
Carbon dioxide is a linear, symmetrical molecule (O=C=O, bond angle 180°), so the two bond dipoles cancel and the molecule has no overall dipole.

Water is V-shaped (bent) because of the two lone pairs of electrons on the oxygen atom, so the dipoles of the two O–H bonds do not cancel and water is a polar molecule.

(c) Pure water ionises to form H_3O^+ and OH^- ions, although only to a very small extent.
Draw the dot-and-cross diagrams of these ions.
Use dots (•) for the hydrogen electrons and crosses (×) for the oxygen electrons.

(2)





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(d) An equation for the ionisation of water is

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

The expression for the ionic product of water is

$$K_w = [H^+(aq)]\,[OH^-(aq)]$$

The value of K_w at 310 K is $2.40 \times 10^{-14}\,mol^2\,dm^{-6}$

(i) Calculate the pH of water at 310 K.

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

(2)

$[H^+] = \sqrt{K_w} = \sqrt{(2.40 \times 10^{-14})} = 1.549 \times 10^{-7}$ mol dm^{-3}

$pH = -\log(1.549 \times 10^{-7}) = 6.81$

$pH = 6.81$

(ii) Predict, with a reason, whether water is acidic, alkaline or neutral at 310 K.

(2)

Neutral, because $[H^+(aq)] = [OH^-(aq)]$: water ionises to give equal amounts of H^+ and OH^- ions, even though the pH is below 7.

(iii) Predict, with a reason, the sign of the enthalpy change for the ionisation of water.

(1)

Positive (ΔH is +), because K_w increases as the temperature increases (from 1.00×10^{-14} at 298 K to 2.40×10^{-14} at 310 K), so the equilibrium shifts to the right and there is more ionisation at the higher temperature.

(Total for Question 6 = 15 marks)

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*7 Colour is often used in chemistry to identify substances.

Compare and contrast the origin of the colour of a copper(II) complex with the origin of the colour of the copper(II) ion in a flame test.

You do not need to state any specific colours.

(6)

In both the copper(II) complex and the flame test, the colour arises from electrons moving between energy levels, and the colour depends on the size of the energy gap (the difference in energy levels), since $\Delta E = hf$.

Flame test: heat energy from the flame promotes (excites) electrons in the copper(II) ions to higher energy levels.

The excited electrons then return (fall) to lower energy levels and release the energy as light. The frequency of this light corresponds to the energy difference between the levels, so the colour seen is the light that is emitted.

Copper(II) complex: the Cu^{2+} ion has a partially filled 3d subshell. When ligands form dative bonds to the Cu^{2+} ion, the d orbitals are split into two groups with different energies by the ligands.

Electrons in the lower energy d orbitals absorb energy from visible light and are promoted to the higher energy d orbitals.

The frequency of the light absorbed corresponds to the energy gap ΔE between the split d orbitals. The remaining frequencies are transmitted or reflected, so the colour seen is the complementary colour, which is the colour not absorbed.

So the main contrast is that in the flame test the energy for promotion comes from heat and the colour seen is the light emitted as electrons fall back, whereas in the complex the energy comes from visible light and the colour seen is the light that is not absorbed.

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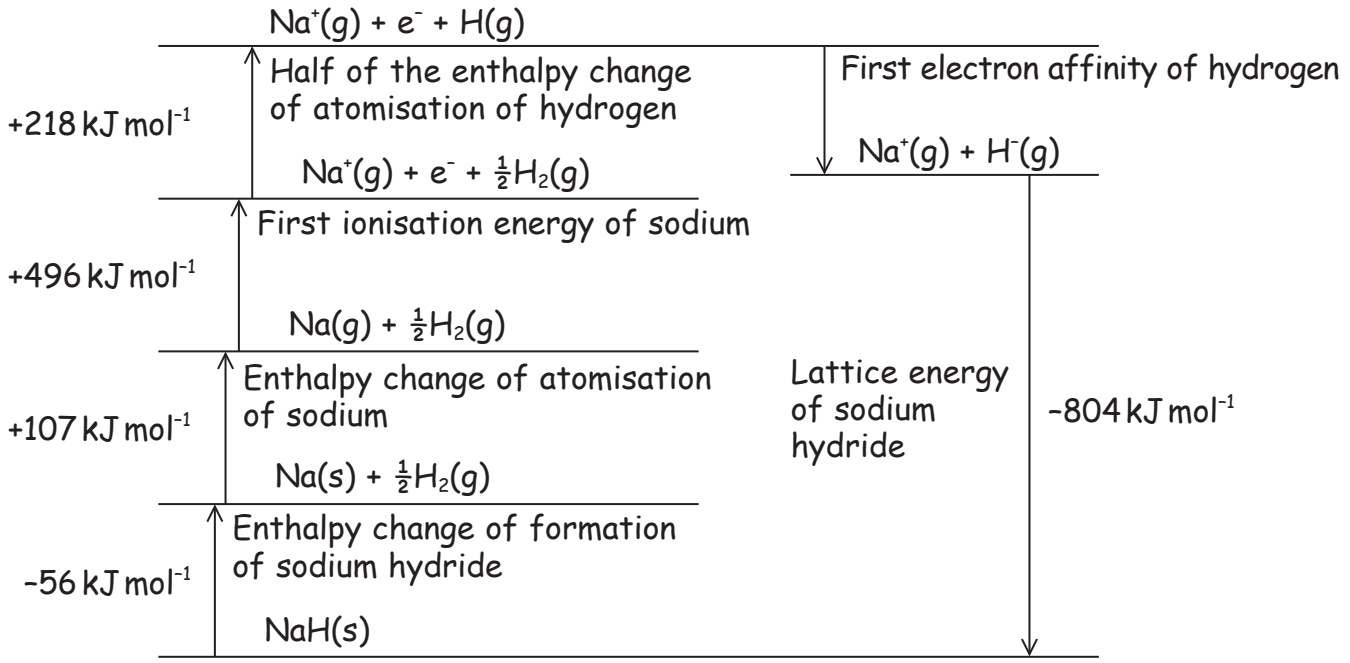
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(Total for Question 7 = 6 marks)

8 Sodium hydride, NaH, can be used to generate hydrogen for fuel cells.

(a) In order to calculate the first electron affinity of hydrogen, a student was asked to draw a Born-Haber cycle for sodium hydride.

The cycle had **two** errors but the numerical data were correct.



(i) Identify and correct the **two** errors in this Born-Haber cycle.

(2)

Error 1: the arrow for the enthalpy change of formation of sodium hydride should point downwards, from Na(s) + ½H₂(g) to NaH(s).

Error 2: the word 'half' should be deleted, as +218 kJ mol⁻¹ is the enthalpy change of atomisation of hydrogen.

(ii) Calculate the first electron affinity, in kJ mol^{-1}, of hydrogen, using the values given in the cycle.

(1)

$EA_1 = -56 - (+107 + 496 + 218) - (-804)$

$= -56 - 821 + 804$

First electron affinity of hydrogen = −73 kJ mol⁻¹

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(b) The equation for the formation of sodium hydride is

$$Na(s) + \tfrac{1}{2}H_2(g) \rightarrow NaH(s) \qquad \Delta_f H^{\ominus} = -56\,kJ\,mol^{-1}$$

The standard entropy change of the system, $\Delta S^{\ominus}_{system}$, for this reaction is $-76.5\,J\,K^{-1}\,mol^{-1}$.

(i) Deduce the feasibility of this reaction at 298 K by calculating the free energy change, ΔG.

(2)

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

$\Delta G = -56 - (298 \times -76.5 \div 1000)$

$\Delta G = -56 + 22.8 = -33.2\ kJ\ mol^{-1}$

ΔG is negative (less than 0), so the reaction is feasible at 298 K.

(ii) Calculate the temperature at which $\Delta G = 0$.

(1)

When $\Delta G = 0$, $T = \Delta H \div \Delta S_{system}$

$T = -56 \div (-0.0765) = 732\ K$

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(c) The sodium hydride is crushed in the presence of water to release the hydrogen gas for a fuel cell.

The overall equation for the reaction occurring in the fuel cell is

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)$$

In an alkaline fuel cell the oxidation half-equation is

$$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$$

Deduce the reduction half-equation for the alkaline fuel cell.

State symbols are not required.

(1)

$\frac{1}{2}O_2 + H_2O + 2e^- \rightarrow 2OH^-$

(d) Lattice energies provide an indication of ionic bond strength.

Which are the lattice energies of the hydrides NaH, KH and MgH_2?

(1)

		Lattice energy / kJ mol^{-1}		
		Sodium hydride, NaH	Potassium hydride, KH	Magnesium hydride, MgH_2
☒	**A**	–804	–711	–1018
☒	**B**	–804	–711	–2718
☒	**C**	–804	–911	–1018
☒	**D**	–804	–911	–2718

(Total for Question 8 = 8 marks)

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9 This is a question about buffer solutions.

(a) A buffer solution always

(1)

☒ **A** keeps the pH less than 7.

☒ **B** contains equimolar amounts of acid and its conjugate base.

☒ **C** keeps the pH constant if small quantities of acid or base are added.

☒ **D** resists changes in pH if small quantities of acid or base are added.

(b) A buffer solution with a pH of 3.90 is required.
Calculate the **mass**, in grams, of sodium ethanoate that should be added to 50.0 cm^3 of an ethanoic acid solution of concentration 0.800 mol dm^{-3} to form this buffer solution.

Give your answer to an appropriate number of significant figures.

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

(5)

$[H^+] = 10^{-3.90} = 1.259 \times 10^{-4}$ mol dm^{-3}

$K_a = [H^+][CH_3COO^-] \div [CH_3COOH]$

$[CH_3COO^-] = (1.74 \times 10^{-5} \times 0.800) \div (1.259 \times 10^{-4})$

$= 0.1106$ mol dm^{-3}

moles of $CH_3COONa = 0.1106 \times 50.0 \div 1000 = 5.529 \times 10^{-3}$ mol

M_r of $CH_3COONa = 82.0$

mass $= 5.529 \times 10^{-3} \times 82.0 = 0.4533$ g

Mass of sodium ethanoate = 0.453 g

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(c) One of the systems controlling the pH of blood is the carbonic acid-hydrogencarbonate buffer system.

$$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$$

Explain how this buffer system helps to control the pH of blood when extra carbon dioxide is present due to strenuous exercise.

(3)

The extra carbon dioxide dissolves in the blood and reacts with water to form carbonic acid, $CO_2 + H_2O \rightarrow H_2CO_3$, so the concentration of H_2CO_3 increases.

The equilibrium $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ shifts to the right, as the carbonic acid partially dissociates to produce more H^+ ions.

However, there is a large reservoir of hydrogencarbonate ions, HCO_3^-, which combine with the extra H^+ ions (the high concentration of HCO_3^- suppresses the ionisation of carbonic acid), so the pH of the blood stays almost constant.

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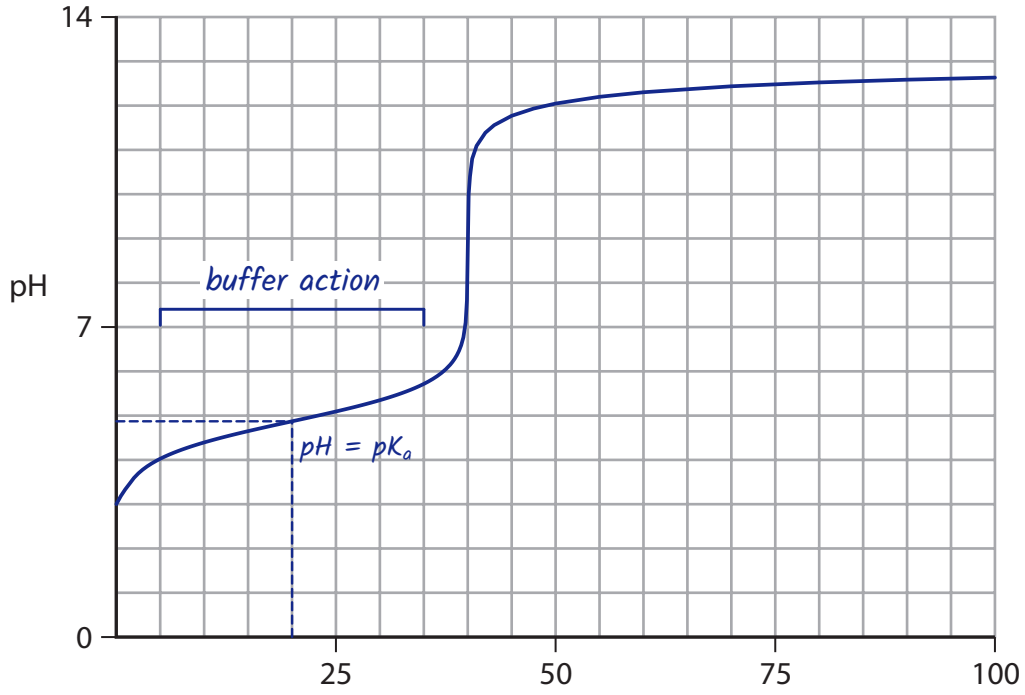
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(d) A weak acid-strong base titration curve can be used to demonstrate buffer action.

(i) Draw a titration curve for the addition of
100 cm^3 of sodium hydroxide solution of concentration 0.100 mol dm^{-3} to
40.0 cm^3 of propanoic acid solution of concentration 0.100 mol dm^{-3}
which has a pH of 3.0.

Show the part of the curve that demonstrates buffer action.

(4)



(ii) Describe, without calculation, how you would use your curve to determine the value of K_a for propanoic acid.

(2)

Find the equivalence point volume (40 cm^3, the middle of the vertical section) and halve it to find the half-neutralisation point, 20 cm^3.

Read the pH at the half-neutralisation point from the curve: here [HA] = [A^-], so pH = pK_a.

Then $K_a = 10^{-pH}$ (the inverse log of minus the pH).

(Total for Question 9 = 15 marks)

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10 This question is about equilibrium systems.

(a) Sulfur dioxide and oxygen form an equilibrium with sulfur trioxide.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

The composition of an equilibrium mixture at 698 K and a total pressure of 2.40 atm is shown in the table.

Substance	$SO_2(g)$	$O_2(g)$	$SO_3(g)$
Number of moles /mol	0.0160	0.0120	0.772

(i) Calculate the value of K_p at this temperature.

Include units, if appropriate.

(5)

total moles = 0.0160 + 0.0120 + 0.772 = 0.800 mol

mole fractions: SO_2 = 0.0160 ÷ 0.800 = 0.0200

O_2 = 0.0120 ÷ 0.800 = 0.0150 SO_3 = 0.772 ÷ 0.800 = 0.965

partial pressures: $p(SO_2)$ = 0.0200 × 2.40 = 0.0480 atm

$p(O_2)$ = 0.0150 × 2.40 = 0.0360 atm $p(SO_3)$ = 0.965 × 2.40 = 2.316 atm

$K_p = p(SO_3)^2 \div (p(SO_2)^2 \times p(O_2))$

$K_p = 2.316^2 \div (0.0480^2 \times 0.0360) = 6.47 \times 10^4$

$K_p = 6.47 \times 10^4$ atm^{-1}

(ii) Calculate the number of sulfur dioxide molecules present in this equilibrium mixture.

(1)

$N = 0.0160 \times 6.02 \times 10^{23}$

$N = 9.63 \times 10^{21}$

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(iii) Deduce, by referring to K_p, how the number of sulfur dioxide molecules will change if more oxygen is added to the equilibrium mixture.

(2)

K_p stays the same, because only temperature changes the value of K_p.

Adding oxygen increases $p(O_2)$ in the denominator, so $p(SO_2)$ must decrease to keep K_p constant: the equilibrium shifts to the right and the number of sulfur dioxide molecules decreases.

(b) An equilibrium exists in aqueous solution between the chromate(VI) ions and the dichromate(VI) ions.

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$$

Explain any change in the position of equilibrium if a few drops of sodium hydroxide solution are added to this equilibrium system.

(2)

The position of equilibrium shifts to the left (towards the chromate(VI) ions).

This is because the hydroxide ions react with (neutralise) the H^+ ions to form water, removing H^+ from the equilibrium. The equilibrium moves to the left to replace some of the H^+ ions.

(c) The equilibrium for the reaction between hydrogen gas and an oxide of iron is

$$Fe_3O_4(s) + 4H_2(g) \rightleftharpoons 3Fe(s) + 4H_2O(g)$$

The K_c expression for this equilibrium is

(1)

☒ **A** $K_c = \dfrac{[Fe] \times [H_2O]}{[Fe_3O_4] \times [H_2]}$

☒ **B** $K_c = \dfrac{[Fe]^3 \times [H_2O]^4}{[Fe_3O_4] \times [H_2]^4}$

☒ **C** $K_c = \dfrac{[H_2O]}{[H_2]}$

☒ **D** $K_c = \dfrac{[H_2O]^4}{[H_2]^4}$

(Total for Question 10 = 11 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 darmstadtium 110	[272] Rg roentgenium 111							

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

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
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

* Lanthanide series

* Actinide series

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