Please check the examination details below before entering your candidate information

Candidate surname	Other names
StemUp Model Solution	

Centre Number | Candidate Number

Pearson Edexcel Level 3 GCE

Time 1 hour 45 minutes

Paper reference **9CH0/01**

Chemistry

Advanced

PAPER 1: Advanced Inorganic and Physical Chemistry

You must have:
Scientific calculator, Data Booklet, ruler

Total Marks

Instructions

- Use **black** ink or 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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Q:1/1/1/1/1/1/

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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 atomic structure and the Periodic Table.

(a) Complete the table to show the relative charges and masses for a neutron and an electron.

(2)

Particle	Relative charge	Relative mass
proton	+1	1
neutron	0	1
electron	−1	1/1840

(b) Atomic emission spectroscopy provides evidence for the existence of

(1)

☒ **A** atoms

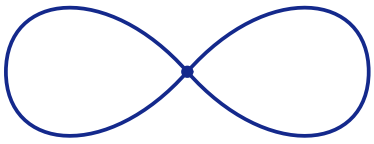
☒ **B** electrons

☒ **C** isotopes

☒ **D** quantum shells

(c) Draw the shape of a p orbital.

(1)



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(d) The melting temperatures of two elements in Period 3 are given in the table.

Element	silicon	chlorine
Melting temperature / K	1683	172

Explain, in terms of the structure and bonding of each element, the difference between these values.

(3)

Silicon has a giant covalent (giant atomic) structure, in which every atom is joined to others by strong covalent bonds.

Chlorine is simple molecular: it is made of Cl_2 molecules held together only by weak London forces (instantaneous dipole to induced dipole attractions).

To melt silicon many strong covalent bonds must be broken, which takes much more energy than overcoming the weak London forces between chlorine molecules, so silicon has the much higher melting temperature.

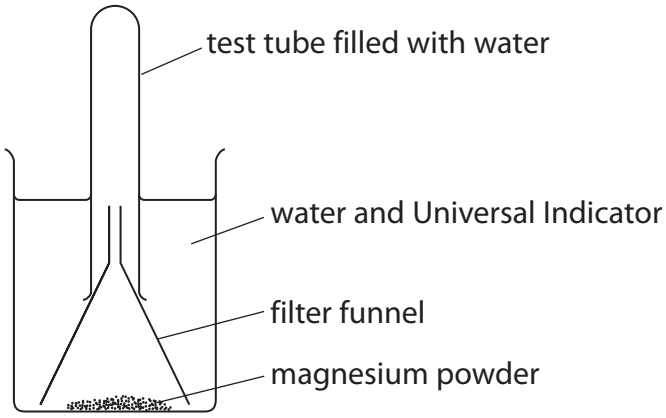
(Total for Question 1 = 7 marks)

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2 This question is about the elements in Group 2 of the Periodic Table.

(a) Magnesium powder is added to a beaker of water containing a few drops of Universal Indicator.

The apparatus is set up as shown and allowed to stand for a few days.



State **two** changes that will be **seen** after a few days.

(2)

The Universal Indicator changes colour from green to blue (as $Mg(OH)_2$ forms).

Gas (hydrogen) collects at the top of the test tube, so the water level in the test tube drops.

(b) Explain how the trend in the reactivity of the Group 2 elements is determined by their electronic configurations.

(3)

Going down Group 2, the outer s electrons are in a higher quantum shell, so they are further from the nucleus.

There is also more shielding by the filled inner shells of electrons, so the attraction between the nucleus and the outer electrons is weaker.

So the first and second ionisation energies decrease down the group: the outer electrons are lost more easily, so the reactivity increases down the group.

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(c) Calcium reacts with chlorine.

$$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$$

Explain, in terms of electrons, why this is a redox reaction.

(2)

Calcium is oxidised because each calcium atom loses two electrons: $Ca \rightarrow Ca^{2+} + 2e^-$

Chlorine is reduced because each chlorine molecule gains two electrons: $Cl_2 + 2e^- \rightarrow 2Cl^-$

(d) An experiment was carried out to determine the molar volume of hydrogen at room temperature.

0.035 g of magnesium was added to excess hydrochloric acid and 32 cm^3 of hydrogen was produced.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Calculate the molar volume of hydrogen from the results of this experiment. Include units in your answer.

(2)

mol Mg = $0.035 \div 24.3 = 1.440 \times 10^{-3}$ mol = mol H_2

molar volume = $32 \div (1.440 \times 10^{-3})$

= 22 200 cm^3 mol^{-1} (= 22.2 dm^3 mol^{-1})

(Total for Question 2 = 9 marks)

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3 This question is about the elements in Group 7 of the Periodic Table and some of their compounds.

(a) What is the colour of iodine in the solid and gas states?

(1)

		Colour of iodine solid	Colour of iodine gas
☒	**A**	purple	brown
☒	**B**	purple	purple
☒	**C**	grey/black	brown
☒	**D**	grey/black	purple

(b) Which of these reactions occurs?

(1)

☒ **A** $Br_2(aq) + 2NaCl(aq) \rightarrow 2NaBr(aq) + Cl_2(aq)$

☒ **B** $Br_2(aq) + 2NaF(aq) \rightarrow 2NaBr(aq) + F_2(aq)$

☒ **C** $Cl_2(aq) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$

☒ **D** $Cl_2(aq) + 2NaF(aq) \rightarrow 2NaCl(aq) + F_2(aq)$

(c) The halogens can form halide ions during reactions.

Complete the electronic configuration of the chloride ion.

(1)

$1s^2$ $2s^2 2p^6 3s^2 3p^6$

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(d) The standard electrode potentials for two half-equations involving bromine are given.

$$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq) \qquad E^{\ominus} = +1.09\,V$$

$$2HOBr(aq) + 2H^+(aq) + 2e^- \rightleftharpoons Br_2(aq) + 2H_2O(l) \qquad E^{\ominus} = +1.57\,V$$

(i) Explain why the disproportionation of bromine in water is **not** thermodynamically feasible under standard conditions.
Include the overall equation for the disproportionation and its $E^{\ominus}_{cell}$ value.

(3)

$Br_2(aq) + H_2O(l) \rightarrow HOBr(aq) + HBr(aq)$

$E^{\ominus}_{cell} = E^{\ominus}(reduction) - E^{\ominus}(oxidation)$

$E^{\ominus}_{cell} = +1.09 - (+1.57) = -0.48\ V$

$E^{\ominus}_{cell}$ is negative (less than 0), so the disproportionation of bromine in water is not thermodynamically feasible under standard conditions.

(ii) Bromine disproportionates in water to a small extent at 298 K.

Give a possible reason why this reaction occurs.

(1)

Disproportionation is an equilibrium: although K is very small, there is still a small concentration of HOBr and HBr at equilibrium. Also a large excess of water is used, so concentrations are not 1 mol dm^{-3}.

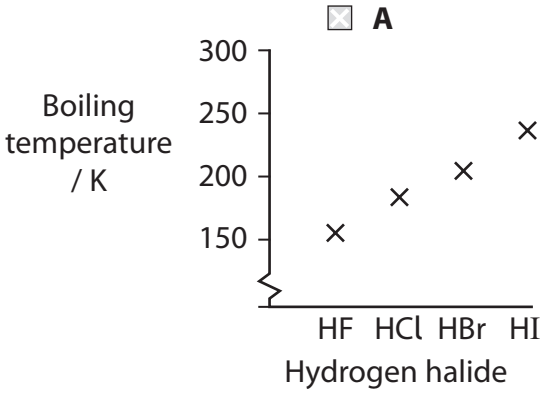
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(e) The hydrogen halides have the general formula HX, where X represents the symbol of the halogen.

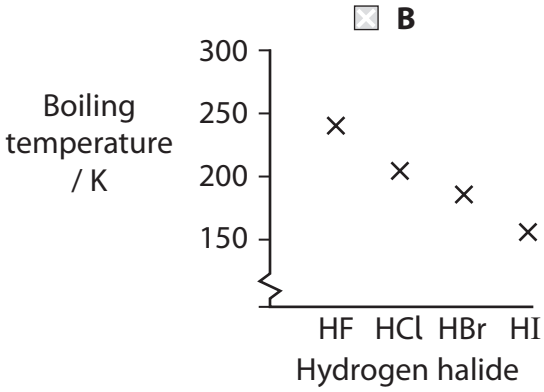
(i) Which diagram shows the trend in the boiling temperatures of the hydrogen halides?

(1)

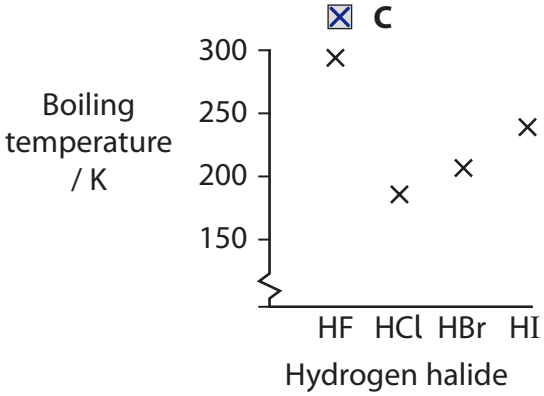
☒ **A**



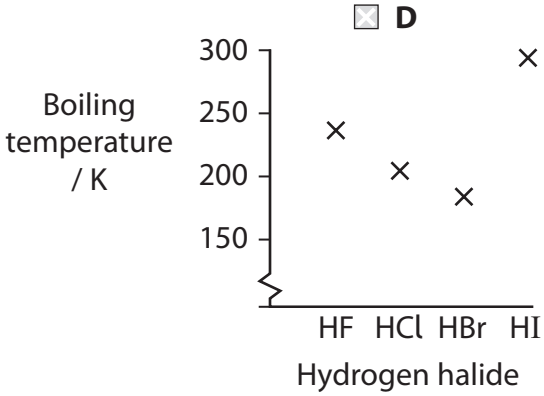
☒ **B**



☒ **C**



☒ **D**



(ii) What type of reaction occurs when ammonia gas reacts with hydrogen chloride gas?

(1)

☒ **A** acid-base

☐ **B** displacement

☐ **C** redox

☐ **D** substitution

(Total for Question 3 = 9 marks)

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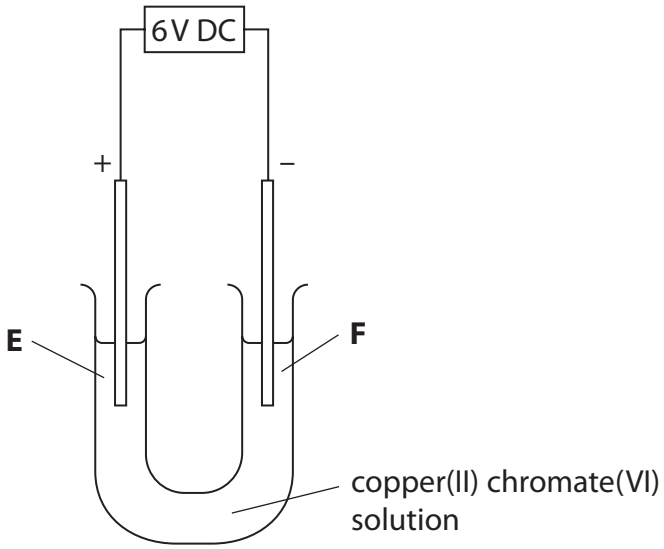
4 This question is about structure and bonding.

(a) Ionic bonding is the strong electrostatic attraction between

(1)

- ☒ **A** anions and cations
- ☐ **B** atoms and delocalised electrons
- ☐ **C** cations and delocalised electrons
- ☐ **D** two nuclei and a shared pair of electrons

(b) An aqueous solution of copper(II) chromate(VI) was electrolysed using the apparatus shown in the diagram.



Deduce the colours of the solutions in regions **E** and **F** after the electrolysis has occurred.

(2)

Colour in region **E** yellow

Colour in region **F** blue

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(c) Some ionic radii are given in the table.

Ion	Ionic radius / nm
Na^+	0.102
Mg^{2+}	0.072
Cl^-	0.180
Br^-	0.195

Deduce the **formula** of the compound, formed from the ions in the table, that has the strongest ionic bonding.

(1)

$MgCl_2$

(d) The names of four substances are given.

Substance	Name
P	copper
Q	iodine
R	silicon(IV) oxide
S	sodium chloride

(i) Which of these substances exists at room temperature as a giant lattice of oppositely charged ions?

(1)

☒ **A** Substance **P**

☒ **B** Substance **Q**

☒ **C** Substance **R**

☒ **D** Substance **S**

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(ii) Which of these substances has a high melting temperature, **and** conducts electricity when solid and when molten?

(1)

☒ **A** Substance **P**

☒ **B** Substance **Q**

☒ **C** Substance **R**

☒ **D** Substance **S**

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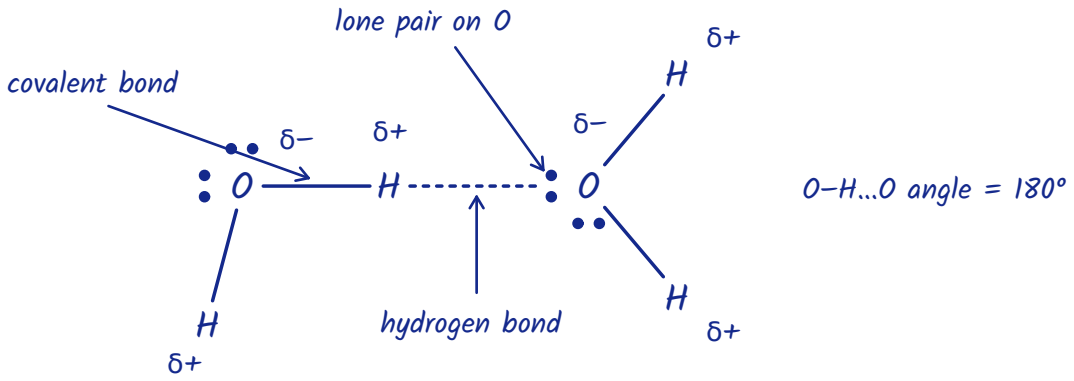
*(e) Water has two significant anomalous properties:

- it has a higher melting temperature than hydrogen sulfide, H_2S, even though it has fewer electrons in its molecules
- the density of ice at 0 °C is less than that of water at 0 °C.

Explain these properties.

You should include a labelled diagram to show the intermolecular forces between two molecules of water.

(6)



Water molecules form hydrogen bonds with each other. Oxygen is much more electronegative than hydrogen, so the O–H bond is polar ($O^{\delta-}$, $H^{\delta+}$). The $H^{\delta+}$ on one molecule is attracted to a lone pair on the $O^{\delta-}$ of another molecule, and the O–H...O arrangement is linear (180°), as shown in the diagram.

H_2S has more electrons per molecule than H_2O, so H_2S has stronger London forces (instantaneous dipole to induced dipole attractions) than water.

However, sulfur is not electronegative enough for H_2S to form hydrogen bonds, so there are no hydrogen bonds between H_2S molecules.

Hydrogen bonds are much stronger than London forces, so much more energy is needed to overcome the intermolecular forces between water molecules. Therefore water has the higher melting temperature even though its molecules have fewer electrons.

In ice at 0 °C, each water molecule is hydrogen bonded to four others, which holds the molecules in a rigid, open hexagonal lattice with large spaces between the molecules

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(the hydrogen bonds are longer than the covalent O–H bonds).

When ice melts, some of the hydrogen bonds break and the open lattice collapses, so the water molecules move closer together and fill the spaces. There are more molecules in the same volume.

Therefore the density of ice at 0 °C is less than that of liquid water at 0 °C.

(Total for Question 4 = 12 marks)

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5 This question is about enthalpy changes and energy changes.

(a) Magnesium carbonate reacts with dilute hydrochloric acid at room temperature.

$$MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$$

When the reaction is carried out in a sealed container with a constant volume, the heat energy change is not the same as the enthalpy change for this reaction.

Give a reason why this is so.

(1)

An enthalpy change is the heat energy change at constant pressure, but carbon dioxide gas is produced in the sealed container of constant volume, so the pressure increases (the system is not at constant pressure).

(b) State what is meant by the standard enthalpy change of formation of aluminium oxide, $Al_2O_3(s)$. Include standard conditions.

(3)

The enthalpy change when one mole of aluminium oxide is formed from its elements in their standard states, under standard conditions of 100 kPa pressure and a stated temperature (usually 298 K).

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

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(c) Use the data in the table to answer the questions.

Enthalpy change	Value / kJ mol^{-1}
Enthalpy change of hydration of K^+	–322
Enthalpy change of hydration of Ca^{2+}	–1650
Enthalpy change of solution of KCl	+17.2
Lattice energy of KCl	–711

(i) Name the two properties of ions that affect the value of their enthalpy change of hydration.

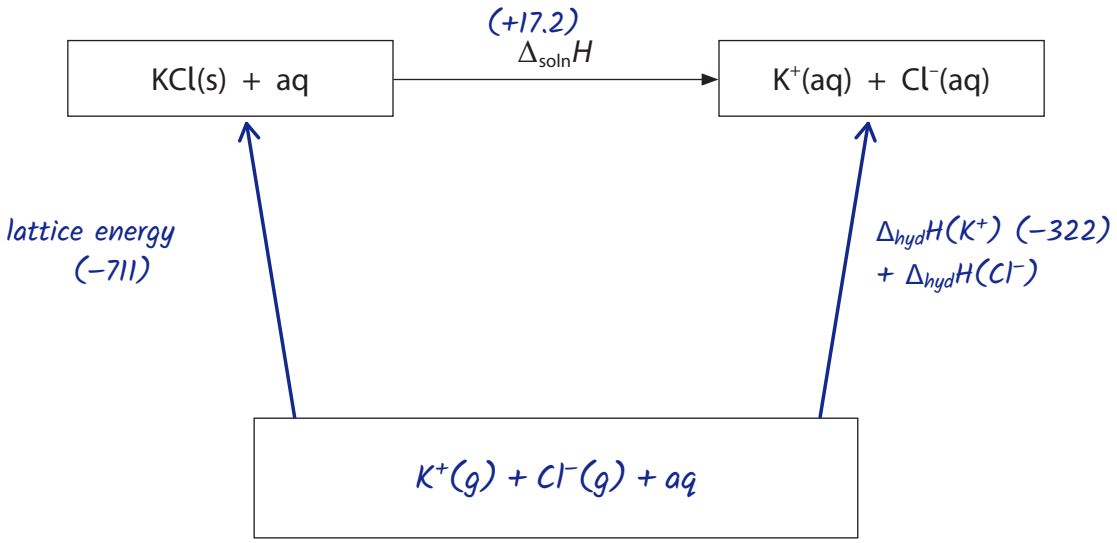
(2)

The ionic radius (size of the ion).

The ionic charge (charge on the ion).

(ii) Calculate the enthalpy change of hydration for chloride ions by completing the energy cycle, including labels, and using the data in the table.

(3)



$\Delta_{hyd}H(Cl^-)$ = lattice energy + $\Delta_{soln}H$ – $\Delta_{hyd}H(K^+)$

= –711 + 17.2 – (–322) = –371.8 kJ mol^{-1}

enthalpy change of hydration for Cl^- ions ..–371.8.. kJ mol^{-1}

(Total for Question 5 = 9 marks)

P 6 7 0 9 3 R A 0 1 5 3 2

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6 This question is about acids and bases.

(a) State what is meant by a Brønsted-Lowry base.

(1)

A proton (H^+ ion) acceptor.

(b) Write the ionic equation for the reaction between magnesium oxide and an acid. State symbols **are** required.

(2)

$MgO(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(l)$

(c) Calculate the concentration of hydrogen ions, in mol dm^{-3}, in a solution with a pH of 9.43

(1)

$[H^+] = 10^{-9.43} = 3.72 \times 10^{-10}$ mol dm^{-3}

(d) The pH of two salt solutions, **J** and **K**, are

solution **J** pH = 5

solution **K** pH = 9

The solutions are equimolar.

Which acids and bases could form the salts in solutions **J** and **K**?

(1)

	Acid and base forming the salt in solution **J**	Acid and base forming the salt in solution **K**
☒ **A**	HCl(aq) and NH_3(aq)	CH_3COOH(aq) and NaOH(aq)
☐ **B**	HCl(aq) and NaOH(aq)	CH_3COOH(aq) and NH_3(aq)
☐ **C**	CH_3COOH(aq) and NaOH(aq)	HCl(aq) and NaOH(aq)
☐ **D**	CH_3COOH(aq) and NH_3(aq)	HCl(aq) and NH_3(aq)

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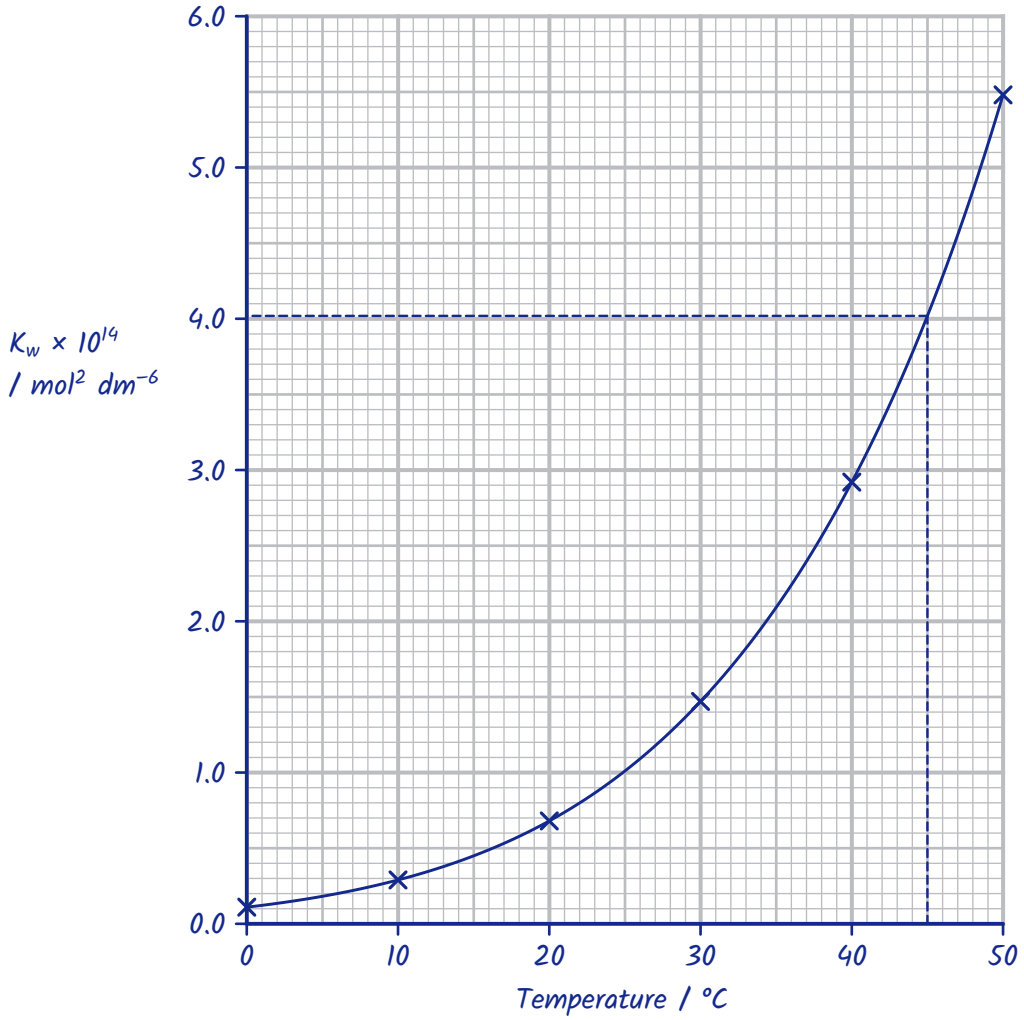
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(e) The ionic product of water, K_w, varies with temperature as shown.

Temperature / °C	K_w / mol^2 dm^{-6}
0	0.11×10^{-14}
10	0.29×10^{-14}
20	0.68×10^{-14}
30	1.47×10^{-14}
40	2.92×10^{-14}
50	5.48×10^{-14}

(i) Determine the value of K_w at 45 °C by plotting a suitable graph.
You must show your working on the graph.

(3)



K_w at 45 °C = 4.0 × 10^{-14} mol^2 dm^{-6}

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(ii) The ionic product of water at 30 °C is 1.47×10^{-14} mol^2 dm^{-6}.

Calculate the pH of water at this temperature.

(3)

$K_w = [H^+][OH^-]$ and in pure water $[H^+] = [OH^-]$

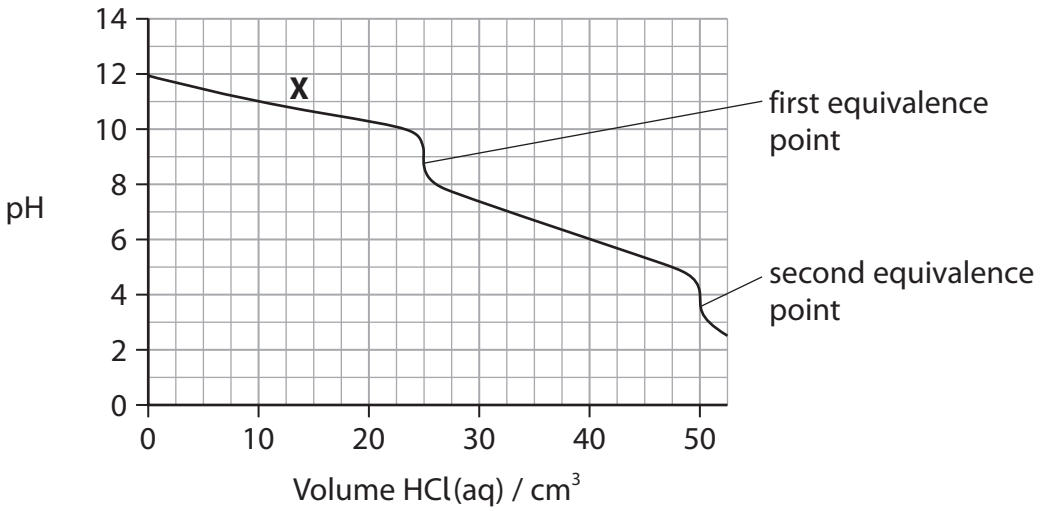
so $K_w = [H^+]^2$

$[H^+] = \sqrt{(1.47 \times 10^{-14})} = 1.212 \times 10^{-7}$ mol dm^{-3}

$pH = -\log(1.212 \times 10^{-7}) = 6.92$

(f) Hydrochloric acid, with a concentration of 0.100 mol dm^{-3}, is added to 25.0 cm^3 of 0.100 mol dm^{-3} aqueous sodium carbonate and the pH is measured.

The titration curve is shown.



The reaction takes place in two steps.
The equation for the reaction taking place in the first step is

$$Na_2CO_3(aq) + HCl(aq) \rightarrow NaHCO_3(aq) + NaCl(aq)$$

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(i) Deduce a suitable indicator to identify the first equivalence point.
Justify your answer using values from the Data Booklet.

(2)

Phenolphthalein.

The pH at the first equivalence point is about 9, which is very close to the pK_{in} of phenolphthalein (9.3), and its pH range of 8.2 to 10.0 lies within the vertical section of the curve at the first equivalence point.

(ii) Write the equation for the reaction taking place at the second equivalence point.
State symbols are not required.

(1)

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

(iii) Explain how the solution at point **X** on the graph can act as a buffer solution.

(3)

At X the solution contains a large reservoir of both carbonate ions, CO_3^{2-}, and hydrogencarbonate ions, HCO_3^- (an acid and its conjugate base).

When a small amount of acid is added, the H^+ ions react with carbonate ions: $CO_3^{2-} + H^+ \rightarrow HCO_3^-$

When a small amount of alkali is added, the OH^- ions react with hydrogencarbonate ions: $HCO_3^- + OH^- \rightarrow CO_3^{2-} + H_2O$

So the ratio of $[CO_3^{2-}]$ to $[HCO_3^-]$ hardly changes and the pH stays almost constant.

(Total for Question 6 = 17 marks)

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7 This question is about chromium and some of its compounds.

(a) The common oxidation numbers of chromium are +2, +3 and +6.

Give a reason, in terms of ionisation energies, why chromium can show variable oxidation numbers.

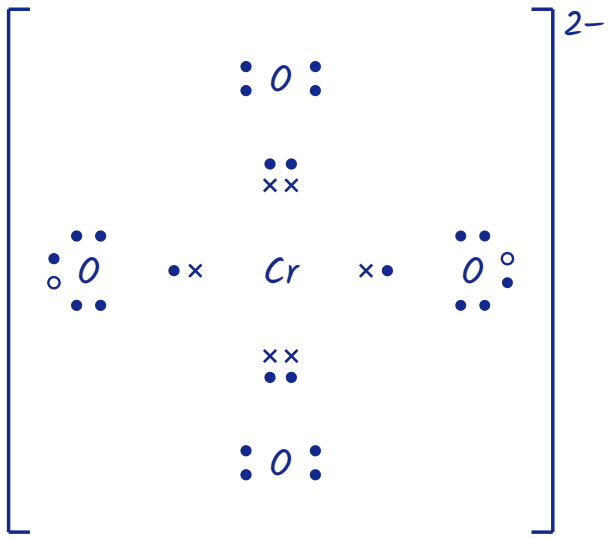
(1)

The successive ionisation energies of chromium increase only gradually (there is no large jump between them), so different numbers of electrons can be removed, giving different oxidation numbers.

(b) The bonding in chromate(VI) ions, CrO_4^{2-}, is similar to that in sulfate(VI) ions, SO_4^{2-}.

Draw a possible dot-and-cross diagram for a chromate(VI) ion.

(2)



× = Cr electron • = O electron ○ = extra electron (2− charge)

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(c) A student added some pieces of zinc to an acidified solution of potassium dichromate(VI).

Some standard electrode potentials are given in the table.

Right-hand electrode system	$E^{\ominus}$ / V
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33

(i) Write the overall equation for the reduction of dichromate(VI) ions to chromium(III) ions by zinc in acid conditions.
State symbols are not required.

(2)

$Cr_2O_7^{2-} + 14H^+ + 3Zn \rightarrow 2Cr^{3+} + 7H_2O + 3Zn^{2+}$

(ii) Calculate $E^{\ominus}_{cell}$ for the reaction in (c)(i).

(1)

$E^{\ominus}_{cell} = +1.33 - (-0.76) = +2.09\ V$

(iii) Predict whether or not a further reduction of chromium(III) ions to chromium(II) ions will occur. Justify your answer.

(1)

Yes. $E^{\ominus}$ for Zn^{2+}/Zn (−0.76 V) is more negative than $E^{\ominus}$ for Cr^{3+}/Cr^{2+} (−0.41 V), so $E^{\ominus}_{cell} = -0.41 - (-0.76) = +0.35$ V, which is positive, so zinc will reduce Cr^{3+} to Cr^{2+}.

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(iv) Aqueous solutions containing chromium(III) ions and chromium(II) ions have different colours.

Explain why these solutions **differ** in colour.
An explanation of the origin of the colours is not required.

(2)

The energy difference between the two sets of split d orbitals is different in Cr^{3+} and Cr^{2+} ions.

So the electrons undergo d to d transitions that absorb a different amount of energy, so a different frequency (wavelength) of visible light is absorbed and a different colour is transmitted.

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(d) An iron nail was analysed using the following outline procedure.

- An iron nail was placed in a beaker and excess dilute sulfuric acid was added.
- After all the iron had reacted to form iron(II) ions, the solution was made up to 1.00 dm^3 in a volumetric flask.
- 25.0 cm^3 portions of the solution were acidified and titrated with potassium dichromate(VI) solution of concentration 0.0167 mol dm^{-3}.

Results

mass of iron nail = 3.54 g

mean titre = 15.50 cm^3

The table shows the percentage by mass of iron in four different brands of nail.

Brand of nail	Percentage by mass of iron
A	92
B	94
C	96
D	98

Potassium dichromate(VI) in acid solution oxidises iron(II) ions as shown in the equation

$$Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

Determine, using the experimental data, the brand of nail that was analysed.

(5)

mol $Cr_2O_7^{2-}$ = 15.50 × 0.0167 ÷ 1000 = 2.5885×10^{-4} mol

mol Fe^{2+} in 25.0 cm^3 = 6 × 2.5885×10^{-4} = 1.5531×10^{-3} mol

mol Fe^{2+} in 1.00 dm^3 = 1.5531×10^{-3} × 1000 ÷ 25.0 = 0.062124 mol

mass of Fe = 0.062124 × 55.8 = 3.4665 g

% Fe = 3.4665 ÷ 3.54 × 100 = 97.9 % (≈ 98 %)

so the nail analysed is brand D

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(Total for Question 7 = 14 marks)

8 This question is about electrode potentials, cells and equilibrium constants.

(a) Chlorine gas can be prepared by the oxidation of chloride ions with manganate(VII) ions in acid solution.

$$MnO_4^-(aq) + 5Cl^-(aq) + 8H^+(aq) \rightleftharpoons Mn^{2+}(aq) + 2\tfrac{1}{2}Cl_2(g) + 4H_2O(l) \qquad E^{\ominus}_{cell} = +0.15\,V$$

During this reaction, each manganate(VII) ion accepts five electrons.

Calculate the equilibrium constant, K, for this reaction at 298 K using the expression

$$\ln K = \frac{nE^{\ominus}_{cell}F}{RT}$$

where n is the number of electrons transferred in the overall equation,
F is the Faraday constant (96 500 C mol^{-1}) and
R is the gas constant (8.31 J mol^{-1} K^{-1}).
Units of K are not required.

(2)

$\ln K = (5 \times 0.15 \times 96\,500) \div (8.31 \times 298)$

$\ln K = 29.23$

$K = e^{29.23} = 4.93 \times 10^{12}$

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(b) A fuel cell produces a voltage from the reaction between a fuel and oxygen.

The reaction occurring at one electrode in a methanol fuel cell is

$$CH_3OH(g) + H_2O(l) \rightarrow HCOOH(aq) + 4H^+(aq) + 4e^-$$

Which reaction occurs at the other electrode?

(1)

☒ **A** $4H^+(aq) + O_2(g) + 4e^- \rightarrow 2H_2O(l)$

☐ **B** $2H_2(g) + 2O_2(g) + 4e^- \rightarrow 4OH^-(aq)$

☐ **C** $4OH^-(aq) \rightarrow 2H_2(g) + 2O_2(g) + 4e^-$

☐ **D** $2H_2O(l) \rightarrow 4H^+(aq) + O_2(g) + 4e^-$

(c) Lead-acid batteries are used as storage cells in some cars.

The electrolyte is sulfuric acid, one electrode is lead and the other is lead(IV) oxide, PbO_2.

As the cell discharges, the lead and the lead(IV) oxide are both converted to solid lead(II) sulfate, $PbSO_4$, and the concentration of the sulfuric acid decreases.

Deduce, using the information given, the two half-equations occurring in the lead-acid battery.
State symbols **are** required.

(3)

$Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$

$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

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(d) When solid lead(II) sulfate is added to aqueous sodium iodide, an equilibrium is established.

$$PbSO_4(s) + 2I^-(aq) \rightleftharpoons PbI_2(s) + SO_4^{2-}(aq)$$

The expression for the equilibrium constant, K_c, for this reaction is

$$K_c = \frac{[SO_4^{2-}(aq)]}{[I^-(aq)]^2}$$

In an experiment, K_c may be determined by adding excess lead(II) sulfate to 25.0 cm^3 of 0.100 mol dm^{-3} sodium iodide.
The volume remains constant at 25.0 cm^3.
The mixture is left to reach equilibrium at room temperature.

Ice-cold water is added to freeze the position of equilibrium and the mixture is then titrated with standard silver nitrate solution.

The whole mixture requires 12.20 cm^3 of 0.0500 mol dm^{-3} silver nitrate solution to react with the aqueous iodide ions at equilibrium.

$$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$$

Calculate the equilibrium concentrations of the sulfate ions and the iodide ions, and hence the value of K_c at room temperature.
Give your answer to an appropriate number of significant figures and include units for K_c, if any.

(7)

initial mol I^- = $25.0 \times 0.100 \div 1000 = 2.50 \times 10^{-3}$ mol

eqm mol I^- = mol Ag^+ = $12.20 \times 0.0500 \div 1000 = 6.10 \times 10^{-4}$ mol

mol I^- reacted = $2.50 \times 10^{-3} - 6.10 \times 10^{-4} = 1.89 \times 10^{-3}$ mol

eqm mol SO_4^{2-} = $1.89 \times 10^{-3} \div 2 = 9.45 \times 10^{-4}$ mol

$[SO_4^{2-}] = 9.45 \times 10^{-4} \times 1000 \div 25.0 = 0.0378$ mol dm^{-3}

$[I^-] = 6.10 \times 10^{-4} \times 1000 \div 25.0 = 0.0244$ mol dm^{-3}

$K_c = 0.0378 \div (0.0244)^2 = 63.5$ dm^3 mol^{-1}

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(Total for Question 8 = 13 marks)

TOTAL FOR PAPER = 90 MARKS

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The Periodic Table of Elements

1	2											3	4	5	6	7	0 (8) (18)
							1.0 H hydrogen 1										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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