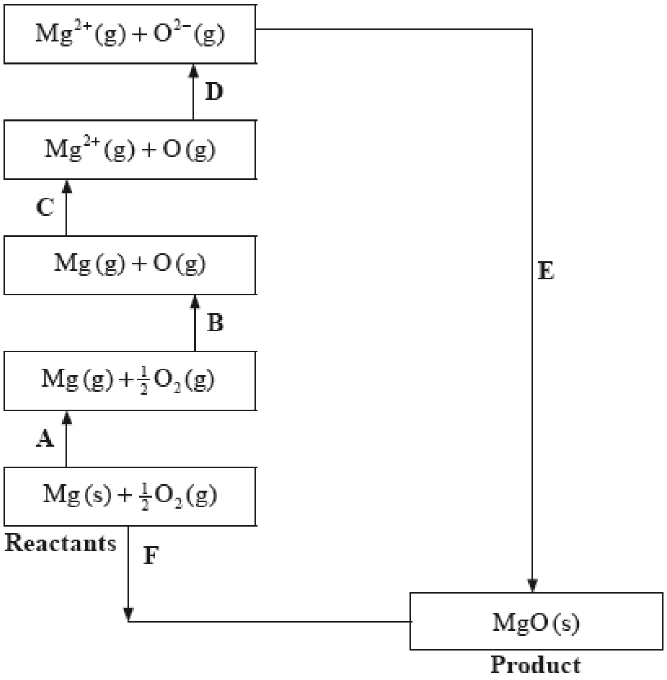
HL Paper 2

The Born-Haber cycle for MgO under standard conditions is shown below.



The values are shown in the table below.

Process	Enthalpy change / kJ mol^{-1}
A	+150
B	+248
C	+736 + (+1450)
D	−142 + (+844)
E	
F	−602

a.i. Identify the processes represented by **A**, **B** and **D** in the cycle. [3]

a.ii. Define the enthalpy change, **F**. [2]

a.iii. Determine the value of the enthalpy change, **E**. [2]

a.iv. Define the enthalpy change **C** for the first value. Explain why the second value is significantly larger than the first. [4]

a.v. The inter-ionic distance between the ions in NaF is very similar to that between the ions in MgO. Suggest with a reason, which compound has the higher lattice enthalpy value. [2]

b.i. The standard enthalpy change of three combustion reactions is given below in kJ. [4]

$$\begin{array}{ll} 2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) & \Delta H^\ominus = -3120 \\ 2H_2(g) + O_2(g) \rightarrow 2H_2O(l) & \Delta H^\ominus = -572 \\ C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) & \Delta H^\ominus = -1411 \end{array}$$

Based on the above information, calculate the standard change in enthalpy, $\Delta H^\ominus$, for the following reaction.

$$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g)$$

b.ii. Predict, stating a reason, whether the sign of $\Delta S^\ominus$ for the above reaction would be positive or negative. [2]

b.iii. Discuss why the above reaction is non-spontaneous at low temperature but becomes spontaneous at high temperatures. [2]

b.iv. Using bond enthalpy values, calculate $\Delta H^\ominus$ for the following reaction. [3]

$$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g)$$

b.v. Suggest with a reason, why the values obtained in parts (b) (i) and (b) (iv) are different. [1]

Table 8 of the Data Booklet shows the atomic and ionic radii of the elements.

a. Describe and explain the trend in atomic radius across period 3. [3]

b. A student formulates the following hypothesis: “If phosphorus were to form a positive ion, P^{3+}, its ionic radius would probably be between 110×10^{-12} m and 212×10^{-12} m.” Evaluate this hypothesis. [2]

Oxidation and reduction can be defined in terms of electron transfer or oxidation numbers.

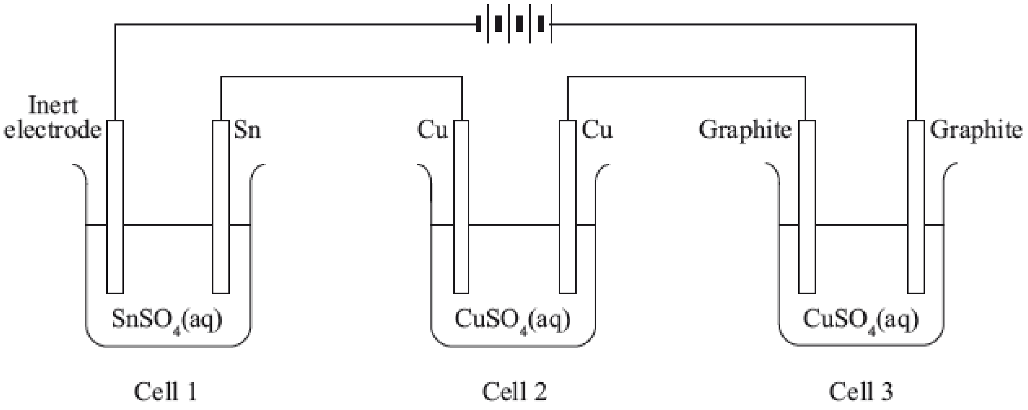
A reactivity series can be experimentally determined by adding the metals W, X, Y and Z to solutions of these metal ions. The following reactions were observed:

$$\begin{array}{l} W^{2+}(aq) + X(s) \rightarrow W(s) + X^{2+}(aq) \\ Y(s) + W^{2+}(aq) \rightarrow Y^{2+}(aq) + W(s) \\ Z^{2+}(aq) + W(s) \rightarrow Z(s) + W^{2+}(aq) \\ Y(s) + X^{2+}(aq) \rightarrow Y^{2+}(aq) + X(s) \end{array}$$

A student carries out the electrolysis of aqueous potassium iodide, KI, using inert electrodes.

Three electrolytic cells were set up in series (one cell after the other), as shown below.

All of the solutions had a concentration of $1.00\ \mathrm{mol\,dm^{-3}}$.



a. Alcohols with the molecular formula C_4H_9OH occur as four structural isomers. Three of the isomers can be oxidized with acidified potassium dichromate solution to form compounds with the molecular formula C_4H_8O. [9]

(i) Deduce the half-equation for the oxidation of the alcohol C_4H_9OH.

(ii) Deduce the overall equation for the redox reaction.

(iii) Two of the isomers with the molecular formula C_4H_9OH can be oxidized further to form compounds with the molecular formula $C_4H_8O_2$. Deduce the structural formulas of these two isomers.

(iv) One isomer cannot be oxidized by acidified potassium dichromate solution.

Deduce its structural formula, state its name and identify it as a primary, secondary or tertiary alcohol.

Name:

Alcohol:

(v) All isomers of the alcohol C_4H_9OH undergo complete combustion. State an equation for the complete combustion of C_4H_9OH.

b. (i) Deduce the order of reactivity of these four metals, from the least to the most reactive. [6]

(ii) A voltaic cell is made by connecting a half-cell of X in $XCl_2(aq)$ to a half-cell of Z in $ZCl_2(aq)$. Deduce the overall equation for the reaction taking place when the cell is operating.

(iii) The standard electrode potential for $Z^{2+}(aq) + 2e^- \rightleftharpoons Z(s)$ is $+0.20\ V$. State which species is oxidized when this half-cell is connected to a standard hydrogen electrode.

(iv) Describe the standard hydrogen electrode including a fully labelled diagram.

c. (i) State the half-equation for the reaction that occurs at each electrode. [4]

Positive electrode (anode):

Negative electrode (cathode):

(ii) Suggest, giving a reason, what would happen if the electrodes were changed to aluminium.

d. (i) Determine the mass of copper produced at one of the electrodes in cell 2 if the tin electrode in cell 1 decreased in mass by 0.034 g. [6]

(ii) Compare the colour and the pH of the solutions in cells 2 and 3 after the current has been flowing for one hour.

(iii) Explain your answer given for part (d) (ii).

Colour:

pH:

Bromine is a member of group 7, the halogens.

Iron is a transition metal.

Freshly prepared iron(II) bromide can be electrolysed both in the liquid state and in aqueous solution.

a.i. Explain the trend in reactivity of the halogens. [3]

a.ii. Deduce, using equations where appropriate, if bromine reacts with sodium chloride solution and with sodium iodide solution. [2]

b.i. Describe the bonding in metals and explain their malleability. [3]

b.ii. List **three** characteristic properties of transition elements. [2]

b.iii. Identify the type of bonding between iron and cyanide in $[\mathrm{Fe(CN)_6}]^{3-}$. [1]

b.iv. Deduce the oxidation number of iron in $[\mathrm{Fe(CN)_6}]^{3-}$. [1]

b.v. Draw the abbreviated orbital diagram for an **iron atom** using the arrow-in-box notation to represent electrons. [1]

b.vi. Draw the abbreviated orbital diagram for the **iron ion in $[\mathrm{Fe(CN)_6}]^{3-}$** using the arrow-in-box notation to represent electrons. [1]

c.i. Describe, using a diagram, the essential components of an electrolytic cell. [3]

c.ii. Describe the **two** ways in which current is conducted in an electrolytic cell. [2]

c.iii. Predict and explain the products of electrolysis of a **dilute** iron(II) bromide solution. [4]

c.iv. Identify another product that is formed if the solution of iron(II) bromide is **concentrated**. [1]

c.v. Explain why this other product is formed. [1]

Calcium nitrate contains both covalent and ionic bonds.

Nitrogen also forms oxides, which are atmospheric pollutants.

a.i. State the formula of both ions present and the nature of the force between these ions. [2]

Ions:

Nature of force:

a.ii. State which atoms are covalently bonded. [1]

b. Bonding in the nitrate ion involves electron delocalization. Explain the meaning of electron delocalization and how it affects the ion. [2]

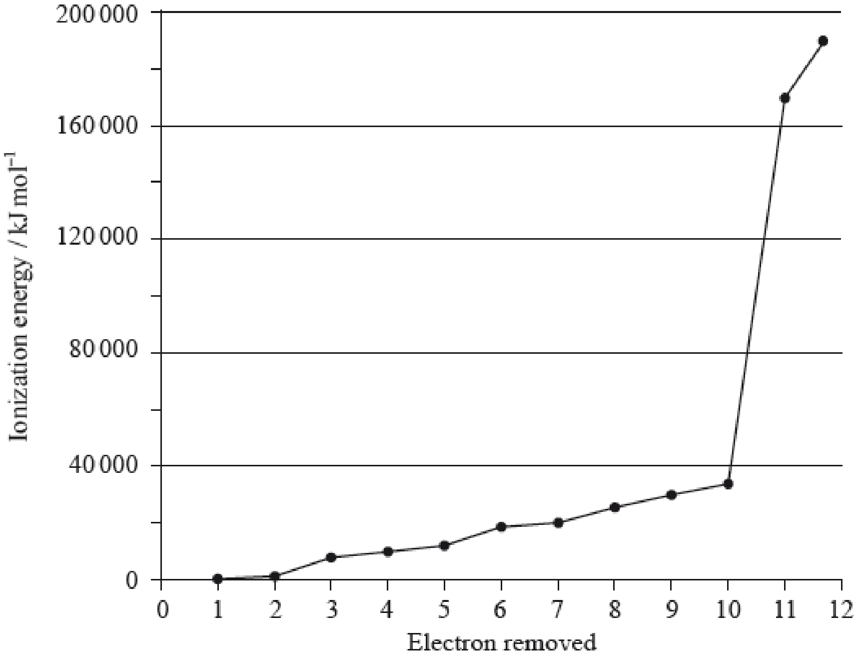
c.i. Outline the source of these oxides. [1]

c.ii. State **one** product formed from their reaction with water. [1]

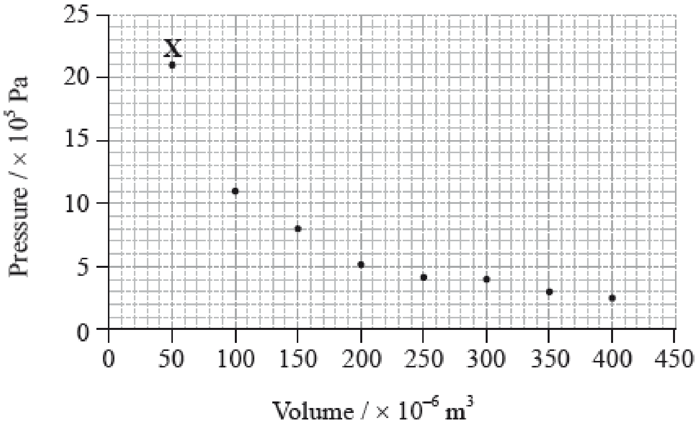
c.iii. State **one** environmental problem caused by these atmospheric pollutants. [1]

A sample of magnesium contains three isotopes: magnesium-24, magnesium-25 and magnesium-26, with abundances of 77.44%, 10.00% and 12.56% respectively.

A graph of the successive ionization energies of magnesium is shown below.



The graph below shows pressure and volume data collected for a sample of carbon dioxide gas at 330 K.



a. (i) Calculate the relative atomic mass of this sample of magnesium correct to **two** decimal places. [4]

(ii) Predict the relative atomic radii of the three magnesium isotopes, giving your reasons.

b. (i) Explain the increase in ionization energy values from the 3rd to the 8th electrons. [3]

(ii) Explain the sharp increase in ionization energy values between the 10th and 11th electrons.

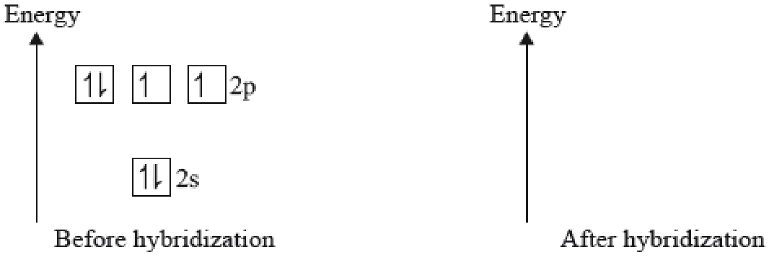
c. (i) Magnesium reacts with oxygen to form an ionic compound, magnesium oxide. Describe how the ions are formed, and the structure and bonding in magnesium oxide. [4]

(ii) Carbon reacts with oxygen to form a covalent compound, carbon dioxide. Describe what is meant by a covalent bond.

(iii) State why magnesium and oxygen form an ionic compound while carbon and oxygen form a covalent compound.

d. (i) Predict the type of hybridization of the carbon and oxygen atoms in $\mathrm{CO_2}$. [7]

(ii) Sketch the orbitals of an oxygen atom in $\mathrm{CO_2}$ on the energy level diagram provided, including the electrons that occupy each orbital.



(iii) Define the term electronegativity.

(iv) Explain why oxygen has a larger electronegativity than carbon.

e. (i) Draw a best-fit curve for the data on the graph. [4]

(ii) Use the data point labelled **X** to determine the amount, in mol, of carbon dioxide gas in the sample.

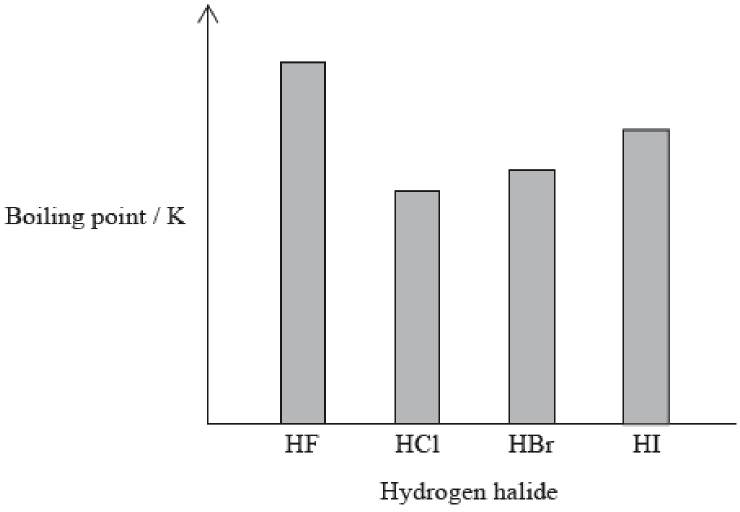
f. (i) Most indicators are weak acids. Describe qualitatively how indicators work. [3]

(ii) Identify a suitable indicator for a titration between a weak acid and a strong base, using Table 16 of the Data Booklet.

The oxides and chlorides of period 3 elements exhibit periodicity.

Chlorine gas, $\mathrm{Cl_2(g)}$, is bubbled through separate solutions of aqueous bromine, $\mathrm{Br_2(aq)}$, and potassium bromide, $\mathrm{KBr(aq)}$.

The hydrogen halides do not show perfect periodicity. A bar chart of boiling points shows that the boiling point of hydrogen fluoride, HF, is much higher than periodic trends would indicate.



Transition metals form complex ions which are usually coloured.

a. (i) State the changes in the acid-base nature of the oxides across period 3 (from Na_2O to Cl_2O_7), including equations for the reactions of Na_2O and SO_3 with water. [7]

(ii) State whether or not molten aluminium chloride, Al_2Cl_6, and molten aluminium oxide, Al_2O_3, conduct electricity. Explain this behaviour in terms of the structure and bonding of the two compounds.

(iii) State the equation for the reaction of Cl_2 with water.

b. (i) Predict any changes that may be observed in each case. [4]

$Br_2(aq)$:

$KBr(aq)$:

(ii) State the half-equations for the reactions that occur.

c. (i) Explain why the boiling point of HF is much higher than the boiling points of the other hydrogen halides. [3]

(ii) Explain the trend in the boiling points of HCl, HBr and HI.

d.i. State the full electron configurations of Cr and Cr^{3+}. [2]

Cr:

Cr^{3+}:

d.ii. Cr^{3+} ions and water molecules bond together to form the complex ion $[\mathrm{Cr(H_2O)_6}]^{3+}$. [3]

Describe how the water acts and how it forms the bond, identifying the acid-base character of the reaction.

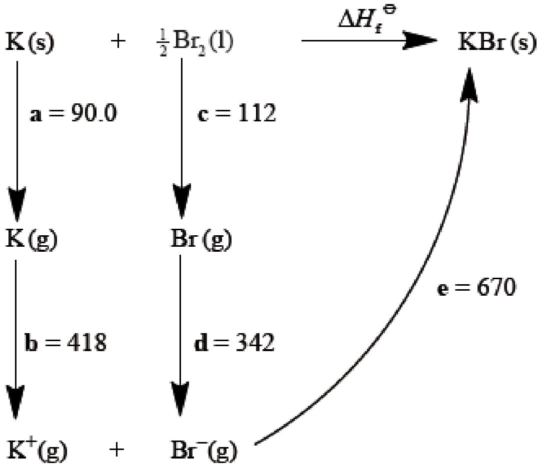
d.iii. Explain why the $[\mathrm{Cr(H_2O)_6}]^{3+}$ ion is coloured. [3]

d.iv. Outline, including a relevant equation, whether the $[\mathrm{Cr(H_2O)_6}]^{3+}$ ion is acidic, basic or neutral. [1]

e. Explain how the number of electrons in the outer main energy level of phosphorus, P, can be determined using the data of successive ionization energies. [2]

Draw the Lewis structures, state the shape and predict the bond angles for the following species.

Consider the following Born-Haber cycle:



The magnitudes for each of the enthalpy changes (**a** to **e**) are given in $\mathrm{kJ\,mol^{-1}}$ but their signs (+ or –) have been omitted.

a.i. $\mathrm{PCl_3}$ [3]

a.ii. $\mathrm{NH_2^-}$ [3]

a.iii. $\mathrm{XeF_4}$ [3]

b.i. State the names for the enthalpy changes **c** and **d**. [2]

b.ii. Deduce which **two** of the enthalpy changes **a** to **e** have negative signs. [1]

b.iii.Determine the value for the enthalpy of formation of potassium bromide. [2]

b.iv.Explain why the quantitative value for the lattice enthalpy of calcium bromide is larger than the value for the lattice enthalpy of potassium bromide. [2]

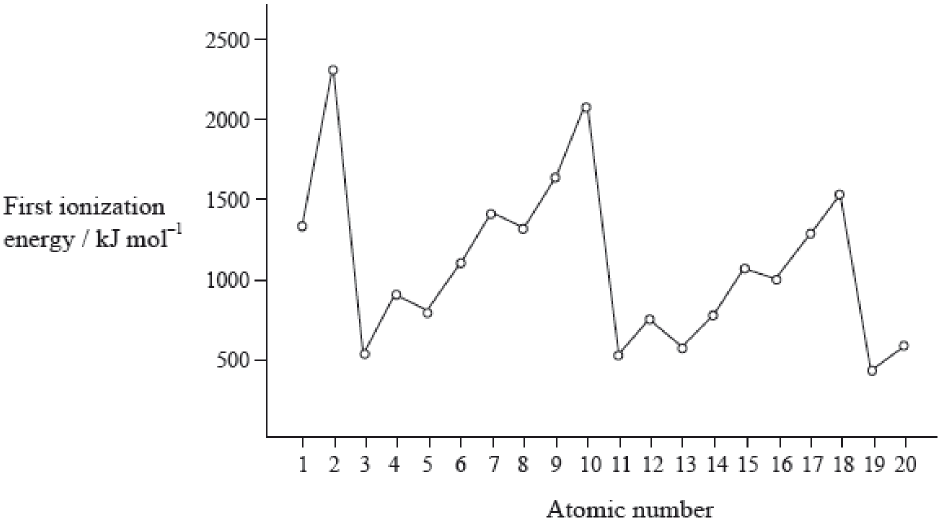
c.i. Compare the formation of a sigma (σ) and a pi (π) bond between two carbon atoms in a molecule. [2]

c.ii.Identify how many sigma and pi bonds are present in propene, C_3H_6. [2]

c.iii.Deduce all the bond angles present in propene. [2]

c.iv.Explain how the concept of hybridization can be used to explain the bonding in the triple bond present in propyne. [3]

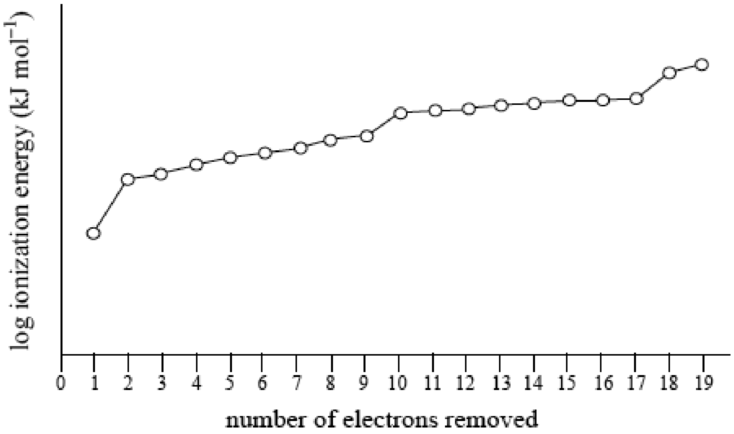
The graph of the first ionization energy plotted against atomic number for the first twenty elements shows periodicity.



b.ii.Explain how information from this graph provides evidence for the existence of main energy levels and sub-levels within atoms. [4]

b.iii.State what is meant by the term *second ionization energy*. [1]

b.iv.Sketch and explain the shape of the graph obtained for the successive ionization energies of potassium using a logarithmic scale for ionization energy on the *y*-axis against number of electrons removed on the *x*-axis. [4]



c.i. State the **full** electronic configurations of copper, Cu, and the copper(I) ion, Cu^{+}. [2]

c.ii. Explain why copper(II) compounds in aqueous solution are coloured whereas scandium(III) compounds in aqueous solution are colourless. [2]

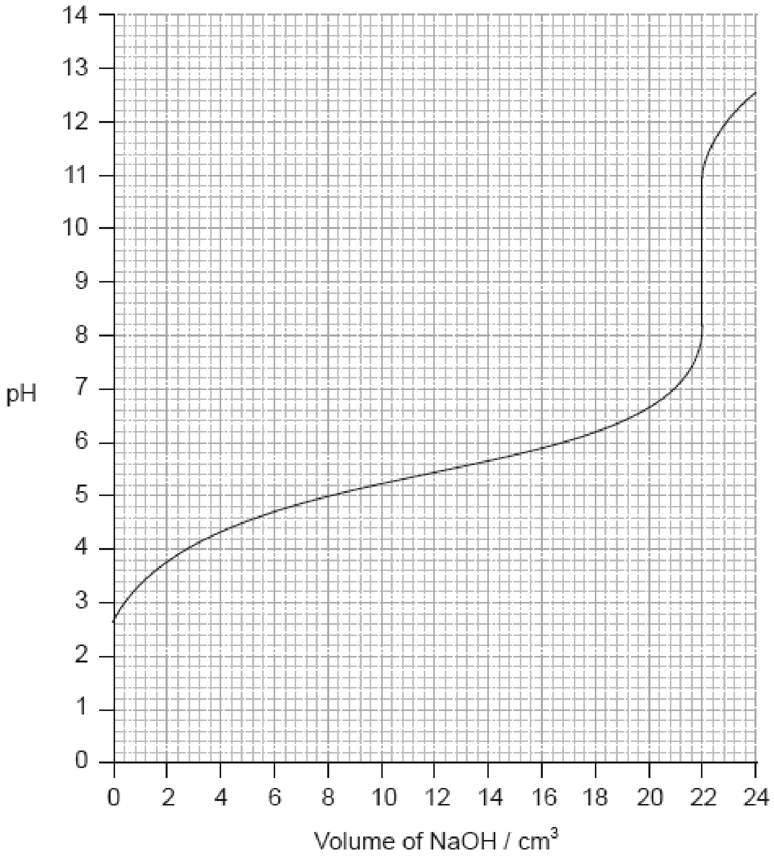
Acids can be described as strong or weak.

a. (i) Outline the difference in dissociation between strong and weak acids of the same concentration. [4]

(ii) Describe **three** tests that can be carried out in the laboratory, and the expected results, to distinguish between $0.10\ mol\,dm^{-3}\ HCl(aq)$ and $0.10\ mol\,dm^{-3}\ CH_3COOH(aq)$.

b. Calculate the pH, using table 15 of the data booklet, of a solution of ethanoic acid made by dissolving 1.40 g of the acid in distilled water to make a $500\ cm^3$ solution. [4]

c.i. Determine the pH at the equivalence point of the titration and the pK_a of an unknown acid using the acid-base titration curve below. [3]



c.ii. Identify, using table 16 of the data booklet, a suitable indicator to show the end-point of this titration. [1]

c.iii Describe how an indicator, that is a weak acid, works. Use Le Chatelier's principle in your answer. [2]

d.i. State the formula of the conjugate base of chloroethanoic acid, $CH_2ClCOOH$. [1]

d.ii. Identify, with a reason, whether chloroethanoic acid is weaker or stronger than ethanoic acid using table 15 of the data booklet. [1]

d.iiiDetermine the pH of the solution resulting when $100\ \mathrm{cm}^3$ of $0.50\ \mathrm{mol\,dm}^{-3}$ $\mathrm{CH_2ClCOOH}$ is mixed with $200\ \mathrm{cm}^3$ of $0.10\ \mathrm{mol\,dm}^{-3}$ NaOH. [4]

e. Describe how chlorine's position in the periodic table is related to its electron arrangement. [2]

f. $\mathrm{SCl_2}$ and $\mathrm{SClF_5}$ are two sulfur chloride type compounds with sulfur having different oxidation states. Predict the name of the shape, the bond angle and polarity of these molecules. [[N/A

The periodic table shows the relationship between electron configuration and the properties of elements and is a valuable tool for making predictions in chemistry.

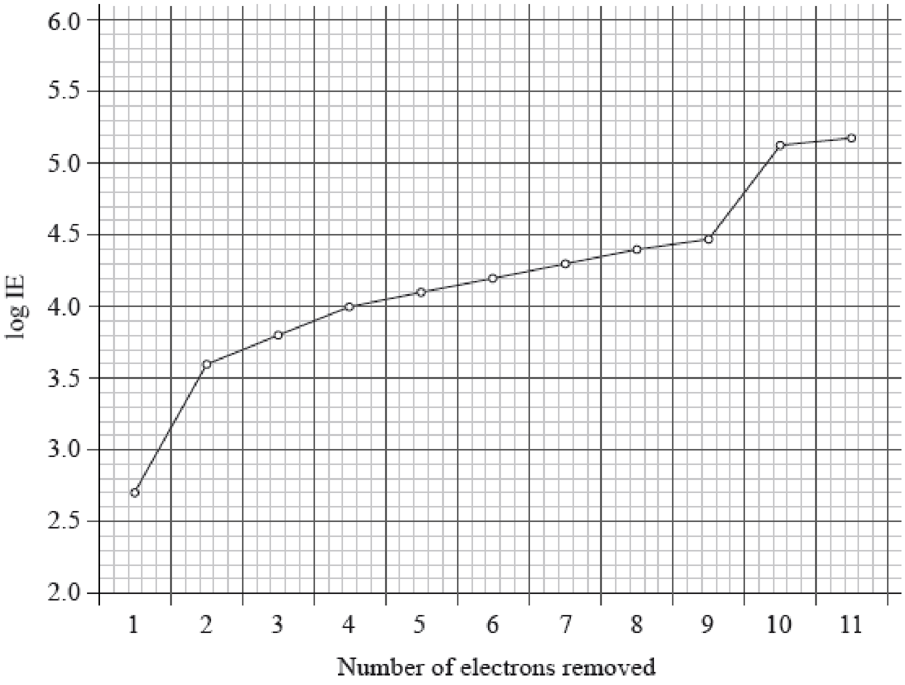
The ten elements in the first-row d-block have characteristic properties and many uses.

b.i. Define the term *electronegativity*. [2]

c. (i) Outline **two** reasons why a sodium ion has a smaller radius than a sodium atom. [4]

(ii) Explain why the ionic radius of $\mathrm{P^{3-}}$ is **greater** than the ionic radius of $\mathrm{Si^{4+}}$.

d. The graph below represents the successive ionization energies of sodium. The vertical axis plots log (ionization energy) instead of ionization energy to allow the data to be represented without using an unreasonably long vertical axis. [4]



State the full electron configuration of sodium and explain how the successive ionization energy data for sodium are related to its electron configuration.

e. (i) Explain why the first ionization energy of aluminium is **lower** than the first ionization energy of magnesium. [4]

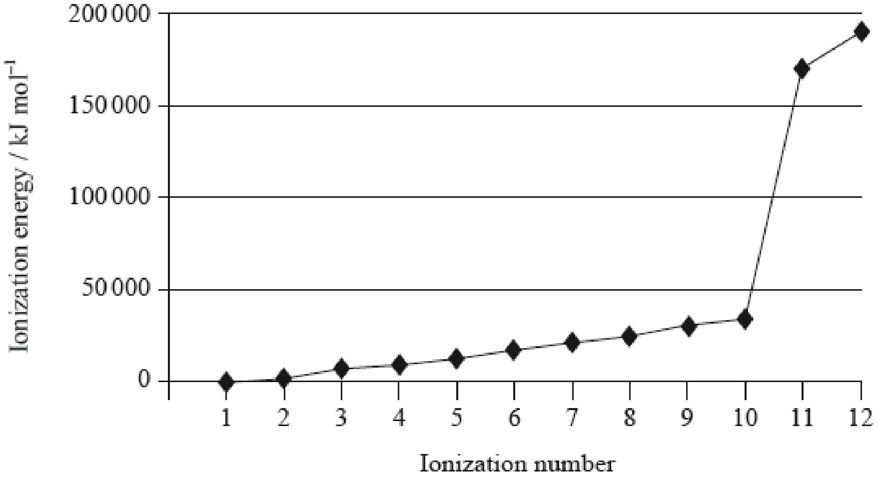
(ii) Explain why the first ionization energy of sulfur is **lower** than the first ionization energy of phosphorus.

f.i. State and explain the type of reaction that takes place between $\mathrm{Fe^{3+}}$ and $\mathrm{H_2O}$ to form $[\mathrm{Fe(H_2O)_6}]^{3+}$ in terms of acid-base theories. [2]

f.ii. Explain why $[\mathrm{Fe(H_2O)_6}]^{3+}$ is coloured. [3]

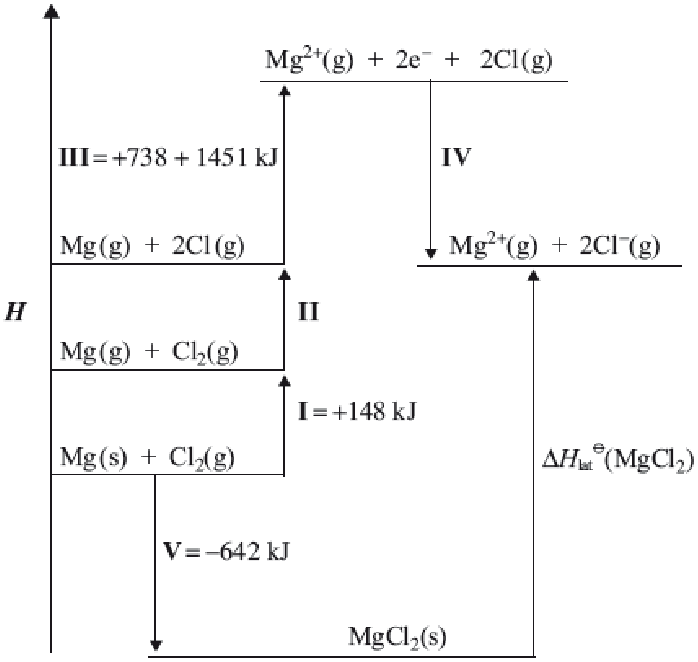
f.iii.Outline the economic significance of the use of a catalyst in the Haber process which is an exothermic reaction. [2]

Magnesium is the eighth most abundant element in the earth's crust. The successive ionization energies of the element are shown below.



Magnesium can be produced from the electrolysis of molten magnesium chloride, $MgCl_2$.

The lattice enthalpy of magnesium chloride can be calculated from the Born-Haber cycle shown below.



a. (i) Define the term *first ionization energy* and state the equation for the first ionization of magnesium. [8]

(ii) Explain the general increase in successive ionization energies of the element.

(iii) Explain the large increase between the tenth and eleventh ionization energies.

b. (i) Explain how molten magnesium chloride conducts an electric current. [5]

(ii) Identify the electrode where oxidation occurs during electrolysis of molten magnesium chloride and state an equation for the half-reaction.

(iii) Explain why magnesium is not formed during the electrolysis of aqueous magnesium chloride solution.

c. (i) Identify the enthalpy changes labelled by **I** and **V** in the cycle. [10]

(ii) Use the ionization energies given in the cycle above and further data from the Data Booklet to calculate a value for the lattice enthalpy of magnesium chloride.

(iii) The theoretically calculated value for the lattice enthalpy of magnesium chloride is +2326 kJ. Explain the difference between the theoretically calculated value and the experimental value.

(iv) The experimental lattice enthalpy of magnesium oxide is given in Table 13 of the Data Booklet. Explain why magnesium oxide has a higher lattice enthalpy than magnesium chloride.

d. (i) State whether aqueous solutions of magnesium oxide and magnesium chloride are acidic, alkaline or neutral. [2]

(ii) State an equation for the reaction between magnesium oxide and water.

This question is about the compounds of some period 3 elements.

a. State the equations for the reactions of sodium oxide with water and phosphorus(V) oxide with water. [2]

b.i. Explain why the melting point of phosphorus(V) oxide is lower than that of sodium oxide in terms of their bonding and structure. [2]

b.ii. Predict whether phosphorus(V) oxide and sodium oxide conduct electricity in their solid and molten states. Complete the boxes with “yes” or “no”. [2]

	Phosphorus(V) oxide	**Sodium oxide**
Solid state	. . . .	. . . .
Molten state	. . . .	. . . .

c. Predict and explain the pH of the following aqueous solutions, using equations to support your answer. [4]

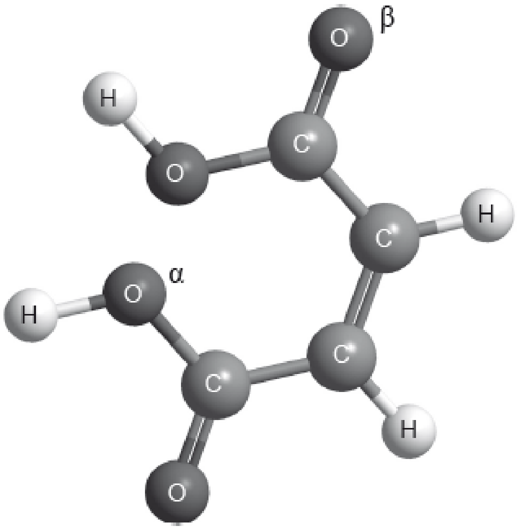
Ammonium chloride, $NH_4Cl(aq)$:

Sodium methanoate, $HCOONa(aq)$:

Consider the structure and bonding in $MgCl_2$ and PCl_3.

Consider the molecules PBr_3 and SF_4.

The structure of *cis*-but-2-ene-1,4-dioic acid is shown below.



a.i. State and explain the electrical conductivities of these two chloride compounds in their liquid state. [3]

a.ii. Suggest, giving your reasons, the approximate pH values of the solutions formed by adding each chloride compound separately to distilled water. [4]

$MgCl_2$

PCl_3

b.i. Identify the acid-base character of the oxides of each of the elements from sodium to chlorine in period 3. [2]

b.ii. State the equations for the separate reactions of sodium oxide and phosphorus(V) oxide with water. [2]

c.i. Deduce the Lewis (electron dot) structure of both molecules. [2]

c.ii. Predict the shapes of the two molecules, giving the Br–P–Br bond angle in PBr_3 and the F–S–F bond angles in SF_4. [4]

PBr_3	SF_4
Shape: . . . . Bond angle:	Shape: . . . . Bond angles: . . . .

c.iii Explain why both PBr_3 and SF_4 are polar. [2]

d.i. Describe the covalent bond between carbon and hydrogen in the molecule above and how it is formed. [2]

d.ii. Deduce the hybridization of the oxygen atoms labelled α and β. [1]

α:

β:

d.iii. Describe sigma (σ) and pi (π) bonds between atoms. [2]

σ bond:

π bond:

d.iv. Identify the number of sigma (σ) and pi (π) bonds present in a molecule of *cis*-but-2-ene-1,4-dioic acid. [1]

Antimony, Sb, forms a fluoride, $\mathrm{SbF_5}$.

The equilibrium that occurs when antimony(V) fluoride is dissolved in liquid hydrogen fluoride can be represented by the equation below.

$$\mathrm{SbF_5(s) + 2HF(l) \rightleftharpoons SbF_6^-(sol) + H_2F^+(sol)}$$

Outline how the following factors account for the fact that HCl is a strong acid and HF is a weak acid.

Some students were provided with a $0.100\ \mathrm{mol\,dm^{-3}}$ solution of a monobasic acid, HQ, and given the problem of determining whether HQ was a weak acid or a strong acid.

The second problem set for the students was to determine the acid dissociation constant, K_a, of the acid HQ and its $\mathrm{p}K_\mathrm{a}$.

a. State the element that you would expect to have chemical properties most similar to those of antimony. [1]

b.i. Describe the relationship between $\mathrm{SbF_5}$ and $\mathrm{SbF_6^-}$ in terms of the Lewis theory of acids. [2]

b.ii. Explain the behaviour of HF in terms of the Brønsted–Lowry theory of acids. [2]

c.i. The strength of the hydrogen–halogen bond. [1]

c.ii. The interaction between an undissociated hydrogen halide molecule and a water molecule. [1]

d.i. Neelu and Charles decided to solve the problem by determining the volume of $0.100\ \mathrm{mol\,dm^{-3}}$ sodium hydroxide solution needed to neutralize $25.0\ \mathrm{cm^3}$ of the acid. Outline whether this was a good choice. [2]

d.ii. Identify **one** indicator that could be used when titrating aqueous sodium hydroxide with both a strong acid and a weak acid, and outline the reason for your choice. [2]

Indicator:

Reason:

d.iii. Neelu and Charles decided to compare the volume of sodium hydroxide solution needed with those required by known $0.100\ \mathrm{mol\,dm^{-3}}$ strong and weak acids. Unfortunately they chose sulfuric acid as the strong acid. Outline why this was an unsuitable choice. [1]

d.iv. Francisco and Shamiso decided to measure the pH of the initial solution, HQ, and they found that its pH was 3.7. Deduce, giving a reason, the strength (weak or strong) of the acid HQ. [2]

e.i. Explain how the $\mathrm{p}K_\mathrm{a}$ could be determined from a graph of pH against the volume of $0.100\ \mathrm{mol\,dm^{-3}}$ sodium hydroxide added. [2]

e.ii. Francisco and Shamiso found that the pH of the initial $0.100\ \mathrm{mol\,dm^{-3}}$ solution was 3.7. However, this reading was inaccurate because they forgot to wash the pH probe. Calculate the $\mathrm{p}K_\mathrm{a}$ of HQ using the reading they obtained. [4]

Bleaches in which chlorine is the active ingredient are the most common, although some environmental groups have concerns about their use.

In aqueous chlorine the equilibrium below produces chloric(I) acid (hypochlorous acid), HOCl, the active bleach.

$$\mathrm{Cl_2(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + H^+(aq) + Cl^-(aq)}$$

Aqueous sodium chlorate(I), NaOCl, the most common active ingredient in chlorine based bleaches, oxidizes coloured materials to colourless products while being reduced to the chloride ion. It will also oxidize sulfur dioxide to the sulfate ion.

The standard electrode potential for the reduction of the chlorate(V) ion to the chloride ion is $+1.49\ \mathrm{V}$.

a. (i) Describe the colour change that occurs when aqueous chlorine is added to aqueous sodium bromide. [3]

(ii) Outline, with the help of a chemical equation, why this reaction occurs.

b.i. Chloric(I) acid is a weak acid, but hydrochloric acid is a strong acid. Outline how this is indicated in the equation above. [1]

b.ii. State a balanced equation for the reaction of chloric(I) acid with water. [1]

b.iii. Outline, in terms of the equilibrium in aqueous chlorine, why it is dangerous to use an acidic toilet cleaner in combination with this kind of bleach. [2]

b.iv. Suggest why a covalent molecule, such as chloric(I) acid, is readily soluble in water. [2]

b.v. Partial neutralization of chloric(I) acid creates a buffer solution. Given that the $\mathrm{p}K_\mathrm{a}$ of chloric(I) acid is 7.53, determine the pH of a solution that has $[\mathrm{HOCl}] = 0.100\ \mathrm{mol\,dm^{-3}}$ and $[\mathrm{ClO^-}] = 0.0500\ \mathrm{mol\,dm^{-3}}$. [4]

b.vi. Describe, using HIn to represent the indicator in its acid form, why an indicator changes colour when excess alkali is added. [3]

c. (i) Deduce a balanced equation for the reaction between the chlorate(I) ion and sulfur dioxide from the appropriate half-equations. [6]

(ii) State the initial and final oxidation numbers of both chlorine and sulfur in the final equation.

Element	Initial oxidation number	Final oxidation number
Chlorine		
Sulfur		

d. (i) Define the term *standard electrode potential*. [3]

(ii) Referring to Table 14 of the Data Booklet, deduce, giving a reason, whether the oxidation of the chromium(III) ion to the dichromate(VI) ion by the chlorate(V) ion is energetically feasible.

Carbon and silicon belong to the same group of the periodic table.

d. Describe the delocalization of pi (π) electrons and explain how this can account for the structure and stability of the carbonate ion, $\mathrm{CO_3^{2-}}$. [3]

e. Explain the meaning of the term *hybridization*. State the type of hybridization shown by the carbon atoms in carbon dioxide, diamond, graphite and the carbonate ion. [5]

f.i. Explain the electrical conductivity of molten sodium oxide and liquid sulfur trioxide. [2]

f.ii. Samples of sodium oxide and solid sulfur trioxide are added to separate beakers of water. Deduce the equation for each reaction and predict the electrical conductivity of each of the solutions formed. [3]

The standard electrode potential for a half-cell made from iron metal in a solution of iron(II) ions, $\mathrm{Fe^{2+}(aq)}$, has the value $-0.45\ \mathrm{V}$.

Consider the following table of standard electrode potentials.

	$E^\ominus$ / V
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$H^+(aq) + e^- \rightleftharpoons \frac{1}{2}H_2(g)$	0.00
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$\frac{1}{2}Br_2(l) + e^- \rightleftharpoons Br^-(aq)$	+1.07

From the list above:

An acidified solution of potassium dichromate is often used as an oxidizing agent in organic chemistry. During the oxidation reaction of ethanol to ethanal the dichromate ion is reduced to chromium(III) ions according to the following **unbalanced** half-equation.

$$\mathrm{Cr_2O_7^{2-}(aq) + H^+(aq) + e^- \rightarrow Cr^{3+}(aq) + H_2O(l)}$$

Sodium metal can be obtained by the electrolysis of molten sodium chloride.

a.i. Define *standard electrode potential*. [2]

a.ii. Explain the significance of the minus sign in -0.45 V. [1]

b.i. State the species which is the strongest oxidizing agent. [1]

b.ii. Deduce which species can reduce $Sn^{4+}(aq)$ to $Sn^{2+}(aq)$ but will not reduce $Sn^{2+}(aq)$ to Sn(s) under standard conditions. [1]

b.iii. Deduce which species can reduce $Sn^{2+}(aq)$ to Sn(s) under standard conditions. [1]

c.i. Draw a labelled diagram of a voltaic cell made from an Fe (s) / $Fe^{2+}(aq)$ half-cell connected to an Ag(s) / $Ag^+(aq)$ half-cell operating under standard conditions. In your diagram identify the positive electrode (cathode), the negative electrode (anode) and the direction of electron flow in the external circuit. [5]

c.ii. Deduce the equation for the chemical reaction occurring when the cell in part (c) (i) is operating under standard conditions and calculate the voltage produced by the cell. [2]

d.i. Describe the colour change that will be observed in the reaction. [1]

d.ii. Deduce the oxidation number of chromium in $Cr_2O_7^{2-}$. [1]

d.iii. State the balanced half-equation for the reduction of dichromate ions to chromium(III) ions. [1]

d.iv. Deduce the half-equation for the oxidation of ethanol to ethanal and hence the overall redox equation for the oxidation of ethanol to ethanal by acidified dichromate ions. [3]

d.v. Explain why it is necessary to carry out the reaction under acidic conditions. [1]

d.vi. Identify the organic product formed if excess potassium dichromate is used and the reaction is carried out under reflux. [1]

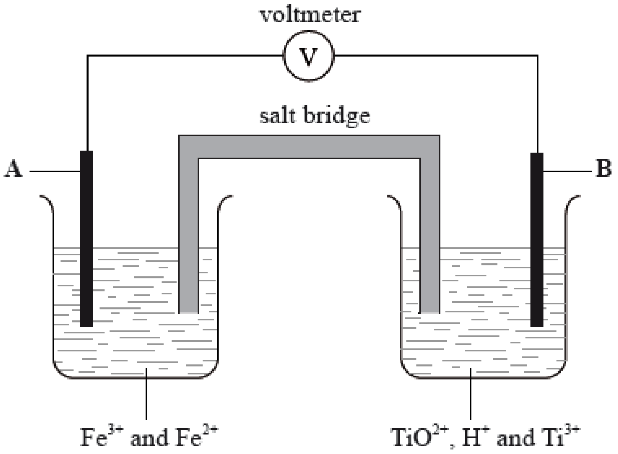
e.i. Explain why it is very difficult to obtain sodium from sodium chloride by any other method. [2]

e.ii. Explain why an aqueous solution of sodium chloride cannot be used to obtain sodium metal by electrolysis. [2]

In acidic solution, ions containing titanium can react according to the half-equation below.

$$\mathrm{TiO^{2+}(aq) + 2H^{+}(aq) + e^{-} \rightleftharpoons Ti^{3+}(aq) + H_2O(l)} \quad E^{\ominus} = -0.06\ \mathrm{V}$$

In the diagram below, **A** and **B** are inert electrodes and, in the aqueous solutions, all ions have a concentration of $1\ \mathrm{mol\,dm^{-3}}$.



Sodium, silicon and sulfur are elements in period 3 of the periodic table that all form oxides.

Although carbon and silicon both belong to group 4 of the periodic table, carbon dioxide and silicon dioxide are different in many ways.

a. Define the term *standard electrode potential*, $E^{\ominus}$. [1]

b.i. State the initial and final oxidation numbers of titanium and hence deduce whether it is oxidized or reduced in this change. [2]

Initial oxidation number	**Final oxidation number**	**Oxidized / reduced**

b.ii. Considering the above equilibrium, predict, giving a reason, how adding more acid would affect the strength of the $\mathrm{TiO^{2+}}$ ion as an oxidizing agent. [2]

c. In the two experiments below, predict whether a reaction would occur and deduce an equation for any reaction that takes place. Refer to Table 14 of the Data Booklet if necessary. [3]

KI(aq) is added to a solution containing $\mathrm{Ti^{3+}(aq)}$ ions:

Zn (s) is added to a solution containing $\mathrm{TiO}^{2+}(\mathrm{aq})$ and $\mathrm{H}^{+}(\mathrm{aq})$ ions:

d.i. Using Table 14 of the Data Booklet, state the balanced half-equation for the reaction that occurs at electrode **A** and whether it involves oxidation or reduction. [2]

d.ii. Calculate the cell potential in V. [1]

d.iii. On the diagram above label with an arrow [1]

- the direction of electron flow in the wire
- the direction in which the positive ions flow in the salt bridge.

e.i. Compare the properties of the three oxides by completing the table below. [3]

	Na_2O	SiO_2	SO_2
Bonding type			
Standard state			
Effect on pH of water			

e.ii. Sulfur dioxide is a significant contributor to acid deposition. Identify a major, man-made source of this pollutant. [1]

e.iii. As well as the oxide above, sodium forms a peroxide that contains the peroxide ion, O_2^{2-}. Draw the Lewis (electron dot) structure of the peroxide ion. [2]

f.i. Describe the differences in the hybridization of these group 4 elements and the precise nature of the bonds that they form with the oxygen atoms. [3]

f.ii. Xenon, although a noble gas, forms an oxide, $\mathrm{XeO_2}$, that has a structure related to that of $\mathrm{SiO_2}$. Compare the geometry around the silicon atoms in $\mathrm{SiO_2}$ with the geometry around the xenon atoms in $\mathrm{XeO_2}$, using the valence shell electron pair repulsion (VSEPR) theory. [3]

Magnesium, a reactive metal found in many common minerals, is also an essential nutrient for both plants and animals.

Successive ionization energies of magnesium are given in the table below.

	First	**Second**	**Third**
Energy required / $kJ\,mol^{-1}$	738	1450	7730

Magnesium metal is mainly used as a component in lightweight alloys, particularly in combination with aluminium and titanium.

Magnesium is usually produced by the electrolysis of molten magnesium chloride.

a. Define the term *first ionization energy*. [2]

b. (i) Explain why the second ionization energy is greater than the first ionization energy. [4]

(ii) Explain why the third ionization energy is much greater than the second ionization energy.

c. Although magnesium is usually found as $\mathrm{Mg^{2+}}$ in its compounds, it is possible to use the Born-Haber cycle to investigate the possibility of $\mathrm{Mg^{+}}$ being able to form stable compounds. [3]

Use the ionization energy data from part (b), along with the other data provided below, to determine the enthalpy change of formation of MgCl(s). Assume that, because $\mathrm{Mg^{+}}$ would be similar in size to $\mathrm{Na^{+}}$, MgCl would have a similar lattice enthalpy to NaCl.

Enthalpy of atomization of Mg	$+146\ \mathrm{kJ\,mol^{-1}}$
Bond enthalpy in $\mathrm{Cl_2}$	$+243\ \mathrm{kJ\,mol^{-1}}$
Electron affinity of Cl	$+349\ \mathrm{kJ\,mol^{-1}}$
Lattice enthalpy of NaCl	$+790\ \mathrm{kJ\,mol^{-1}}$

d. Consider the lattice enthalpies of $\mathrm{MgF_2}$, $\mathrm{MgCl_2}$ and $\mathrm{CaCl_2}$. List these from the most endothermic to the least endothermic and explain your order. [3]

$\text{Most endothermic} \rightarrow \text{Least endothermic}$

e. Magnesium hydroxide, $\mathrm{Mg(OH)_2}$, is only sparingly soluble in water and the equilibrium below exists when excess solid is in contact with a saturated solution. [2]

$$\mathrm{Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^{-}(aq)}$$

Outline how the solubility of magnesium hydroxide will vary with pH.

f. (i) Describe the bonding present in magnesium metal. [4]

(ii) Suggest why magnesium is harder than sodium.

(iii) Outline why alloys are generally less malleable than their component metals.

g. (i) Draw a labelled diagram of a suitable apparatus for the electrolysis. [7]

(ii) State equations for the reactions that take place at the electrodes.

Negative electrode (cathode) reaction:

Positive electrode (anode) reaction:

(iii) When dilute aqueous magnesium chloride is used as the electrolyte, the reactions at both electrodes are different. State equations for the reactions that occur in aqueous solution.

Negative electrode (cathode) reaction:

Positive electrode (anode) reaction:

(iv) Outline why magnesium metal is not produced in the electrolysis of aqueous magnesium chloride.

Nitrogen and silicon belong to different groups in the periodic table.

Draw the Lewis structures, state the shapes and predict the bond angles for the following species.

Consider the molecule $\mathrm{HCONH_2}$.

a.i. Distinguish in terms of electronic structure, between the terms *group* and *period*. [2]

a.ii. State the maximum number of orbitals in the $n = 2$ energy level. [1]

b.i. $\mathrm{SiF_6^{2-}}$ [3]

b.ii $\mathrm{NO_2^+}$ [3]

d. Explain, using diagrams, why $\mathrm{NO_2}$ is a polar molecule but $\mathrm{CO_2}$ is a non-polar molecule. [3]

f.ii. Explain the term *hybridization*. [1]

f.iii. Describe how σ and π bonds form. [2]

f.iv. State the type of hybridization of the carbon and nitrogen atoms in $\mathrm{HCONH_2}$. [2]

Ammonia, $\mathrm{NH_3}$, is a weak base. It has a $\mathrm{p}K_\mathrm{b}$ value of 4.75.

Salts may form neutral, acidic or alkaline solutions when dissolved in water.

a.iii Another weak base is nitrogen trifluoride, $\mathrm{NF_3}$. Explain how $\mathrm{NF_3}$ is able to function as a Lewis base. [1]

a.iv. Calculate the pH of a $1.00 \times 10^{-2}\ \mathrm{mol\,dm^{-3}}$ aqueous solution of ammonia at 298 K. [4]

a.v. $25.0\ \mathrm{cm^3}$ of $1.00 \times 10^{-2}\ \mathrm{mol\,dm^{-3}}$ hydrochloric acid solution is added to $50.0\ \mathrm{cm^3}$ of $1.00 \times 10^{-2}\ \mathrm{mol\,dm^{-3}}$ aqueous ammonia solution. [5]

Calculate the concentrations of both ammonia and ammonium ions in the resulting solution and hence determine the pH of the solution.

a.vi. State what is meant by a buffer solution and explain how the solution in (v), which contains ammonium chloride dissolved in aqueous ammonia, can function as a buffer solution. [3]

b.iii. State the equations for the reactions of sodium oxide, $\mathrm{Na_2O}$, and phosphorus(V)oxide, $\mathrm{P_4O_{10}}$, with water. [2]

Calcium carbide, CaC_2, is an ionic solid.

a. Describe the nature of ionic bonding. [1]

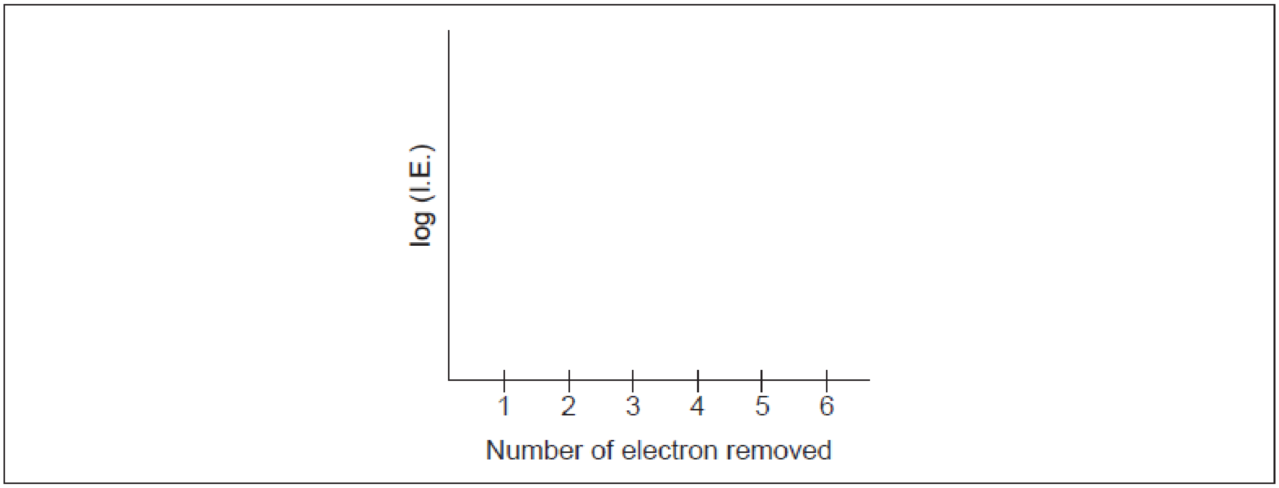
b. Describe how the relative atomic mass of a sample of calcium could be determined from its mass spectrum. [2]

c. When calcium compounds are introduced into a gas flame a red colour is seen; sodium compounds give a yellow flame. Outline the source of the colours and why they are different. [2]

d.i. Suggest **two** reasons why solid calcium has a greater density than solid potassium. [2]

d.ii. Outline why solid calcium is a good conductor of electricity. [1]

e. Sketch a graph of the first six ionization energies of calcium. [2]



f. Calcium carbide reacts with water to form ethyne and calcium hydroxide. [1]

$$CaC_2(s) + H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

Estimate the pH of the resultant solution.

g.i. Describe how sigma (σ) and pi (π) bonds are formed. [2]

sigma (σ):

pi (π):

g.ii.Deduce the number of σ and π bonds in a molecule of ethyne. [1]

sigma (σ):

pi (π):

Trends in physical and chemical properties are useful to chemists.

Cobalt forms the transition metal complex $[Co(NH_3)_4(H_2O)Cl]Br$.

b. Explain why the melting points of the group 1 metals (Li → Cs) decrease down the group whereas the melting points of the group 17 elements (F → I) increase down the group. [3]

Group 1:

Group 17:

d.i. State the shape of the complex ion. [1]

d.ii.Deduce the charge on the complex ion and the oxidation state of cobalt. [2]

Charge on complex ion:

Oxidation state of cobalt:

e. Describe, in terms of acid-base theories, the type of reaction that takes place between the cobalt ion and water to form the complex ion. [2]

Titanium and vanadium are consecutive elements in the first transition metal series.

$\mathrm{TiCl_4}$ reacts with water and the resulting titanium(IV) oxide can be used as a smoke screen.

a. Describe the bonding in metals. [2]

b. Titanium exists as several isotopes. The mass spectrum of a sample of titanium gave the following data: [2]

Mass number	**% abundance**
46	7.98
47	7.32
48	73.99
49	5.46
50	5.25

Calculate the relative atomic mass of titanium to two decimal places.

c. State the number of protons, neutrons and electrons in the ${}^{48}_{22}\mathrm{Ti}$ atom. [1]

Protons:

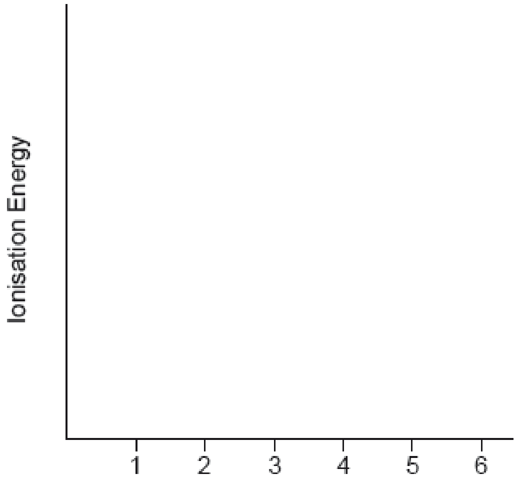
Neutrons:

Electrons:

d.i. State the full electron configuration of the ${}^{48}_{22}\mathrm{Ti}^{2+}$ ion. [1]

d.ii.Suggest why the melting point of vanadium is higher than that of titanium. [1]

d.iii.Sketch a graph of the first six successive ionization energies of vanadium on the axes provided. [1]



d.iv Explain why an aluminium-titanium alloy is harder than pure aluminium. [2]

e. Describe, in terms of the electrons involved, how the bond between a ligand and a central metal ion is formed. [1]

f. Outline why transition metals form coloured compounds. [4]

g.i. State the type of bonding in potassium chloride which melts at 1043 K. [1]

g.ii A chloride of titanium, $\mathrm{TiCl_4}$, melts at 248 K. Suggest why the melting point is so much lower than that of KCl. [1]

h.i. Formulate an equation for this reaction. [2]

h.ii. Suggest one disadvantage of using this smoke in an enclosed space. [1]

Phosphine (IUPAC name phosphane) is a hydride of phosphorus, with the formula PH_3.

a. (i) Draw a Lewis (electron dot) structure of phosphine. [8]

(ii) State the hybridization of the phosphorus atom in phosphine.

(iii) Deduce, giving your reason, whether phosphine would act as a Lewis acid, a Lewis base, or neither.

(iv) Outline whether you expect the bonds in phosphine to be polar or non-polar, giving a brief reason.

(v) Phosphine has a much greater molar mass than ammonia. Explain why phosphine has a significantly lower boiling point than ammonia.

(vi) Ammonia acts as a weak Brønsted–Lowry base when dissolved in water.

$$\mathrm{NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^{\ +}\,(aq) + OH^-\,(aq)}$$

Outline what is meant by the terms “weak” and “Brønsted–Lowry base”.

Weak:

Brønsted–Lowry base:

b. Phosphine is usually prepared by heating white phosphorus, one of the allotropes of phosphorus, with concentrated aqueous sodium hydroxide. The equation for the reaction is: [8]

$$P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$$

(i) The first reagent is written as P_4, not 4P. Describe the difference between P_4 and 4P.

(ii) The ion $H_2PO_2^-$ is amphiprotic. Outline what is meant by amphiprotic, giving the formulas of **both** species it is converted to when it behaves in this manner.

(iii) State the oxidation state of phosphorus in P_4 and $H_2PO_2^-$.

P_4:

$H_2PO_2^-$:

(iv) Oxidation is now defined in terms of change of oxidation number. Explore how earlier definitions of oxidation and reduction may have led to conflicting answers for the conversion of P_4 to $H_2PO_2^-$ and the way in which the use of oxidation numbers has resolved this.

c. 2.478 g of white phosphorus was used to make phosphine according to the equation: [4]

$$P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$$

(i) Calculate the amount, in mol, of white phosphorus used.

(ii) This phosphorus was reacted with 100.0 cm^3 of 5.00 mol dm^{-3} aqueous sodium hydroxide. Deduce, showing your working, which was the limiting reagent.

(iii) Determine the excess amount, in mol, of the other reagent.

(iv) Determine the volume of phosphine, measured in cm^3 at standard temperature and pressure, that was produced.

d. Impurities cause phosphine to ignite spontaneously in air to form an oxide of phosphorus and water. [9]

(i) 200.0 g of air was heated by the energy from the complete combustion of 1.00 mol phosphine. Calculate the temperature rise using section 1 of the data booklet and the data below.

Standard enthalpy of combustion of phosphine, $\Delta H_c^{\ominus} = -750\,kJ\,mol^{-1}$

Specific heat capacity of air = $1.00 Jg^{-1}K^{-1} = 1.00 kJkg^{-1}K^{-1}$

(ii) The oxide formed in the reaction with air contains 43.6% phosphorus by mass. Determine the empirical formula of the oxide, showing your method.

(iii) The molar mass of the oxide is approximately 285 g mol^{-1}. Determine the molecular formula of the oxide.

(iv) State the equation for the reaction of this oxide of phosphorus with water.

(v) Suggest why oxides of phosphorus are not major contributors to acid deposition.

(vi) The levels of sulfur dioxide, a major contributor to acid deposition, can be minimized by either pre-combustion and post-combustion methods. Outline **one** technique of each method.

Pre-combustion:

Post-combustion: