Cambridge

AS Level

Chemistry

CODE: (9701)

Chapter 11 and Chapter 12

Group 2 and Group 17

Chapter 11 – Group 2

Physical properties of Group 2 elements

The elements in Group 2 of the Periodic Table are sometimes referred to as the **alkaline earth metals**.

Here are the electronic configurations of the first five elements in Group 2:

Beryllium (Be)	$1s^2 2s^2$
Magnesium (Mg)	$1s^2 2s^2 2p^6 3s^2$
Calcium (Ca)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Strontium (Sr)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$
Barium (Ba)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2$

The metallic radius, half the distance between nuclei in a giant metallic lattice, is a measure of an atom's size.

Look at the metallic radii of the Group 2 elements, shown in Table 11.1.

Group 2 element	Metallic radius / nm
beryllium (Be)	0.122
magnesium (Mg)	0.160
calcium (Ca)	0.197
strontium (Sr)	0.215
barium (Ba)	0.217

Table 11.1 The metallic radii of the Group 2 elements.

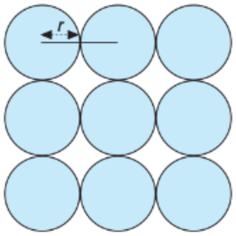


Figure 11.2 The metallic radius gives us a measure of the size of the atoms of metallic elements.

Look at the metallic radii of the Group 2 elements, shown in Table 11.1. There are also general trends in other physical properties, such as melting point and density, shown in Table 11.2 and Figures 11.3 and 11.4.

Group 2 element	Atomic number	Melting point / °C	Density / $g\,cm^{-3}$
beryllium (Be)	4	1280	1.85
magnesium (Mg)	12	650	1.74
calcium (Ca)	20	838	1.55
strontium (Sr)	38	768	2.6
barium (Ba)	56	714	3.5

Table 11.2 The melting points and densities of the Group 2 elements.

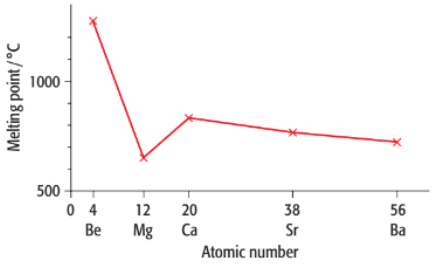


Figure 11.3 Melting points of the Group 2 elements.

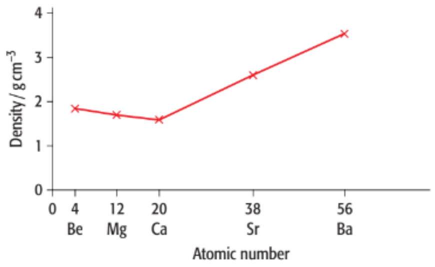
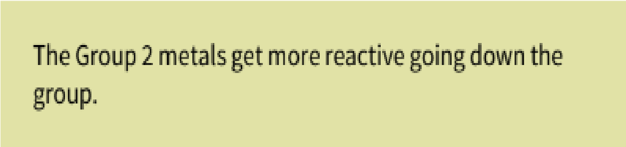


Figure 11.4 Densities of the Group 2 elements.

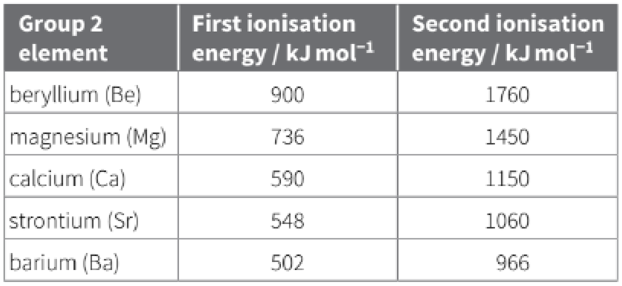
Reactions of Group 2 elements

The Group 2 metals form ionic compounds. When they react, their atoms lose the two electrons from them outermost s subshell and form an ion with the stable electronic configuration of a noble gas. This creates a 2+ ion.

Metals in Group 2 become more reactive as they move down the group, as the removal of outer electron pairs becomes easier. This is due to the greater shielding effect provided by extra inner shells and larger distance from the nucleus.



The Group 2 metals get more reactive going down the group.

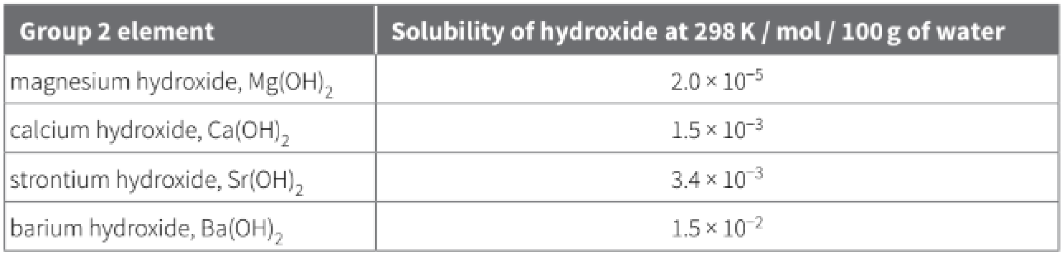


Group 2 element	First ionisation energy / kJ mol^{-1}	Second ionisation energy / kJ mol^{-1}
beryllium (Be)	900	1760
magnesium (Mg)	736	1450
calcium (Ca)	590	1150
strontium (Sr)	548	1060
barium (Ba)	502	966

Table 11.3 The first and second ionisation energies of the Group 2 elements.

The resulting solutions also get more alkaline going down the group. We can explain this by looking at the solubility of the hydroxides formed (see Table 11.4). Their **solubility increases** on going down Group 2.

Not all Group 2 compounds get more soluble on descending the group. For example, their sulfates become **less** soluble – so barium sulfate is much less soluble than magnesium sulfate.



Group 2 element	Solubility of hydroxide at 298 K / mol / 100 g of water
magnesium hydroxide, $Mg(OH)_2$	2.0×10^{-5}
calcium hydroxide, $Ca(OH)_2$	1.5×10^{-3}
strontium hydroxide, $Sr(OH)_2$	3.4×10^{-3}
barium hydroxide, $Ba(OH)_2$	1.5×10^{-2}

Table 11.4 Solubility of the Group 2 hydroxides in water.

Reaction of Group 2 carbonates with water and with dilute acids

The carbonates of magnesium, calcium, strontium and barium are all insoluble in water.

The magnesium sulfate salt formed in the reaction above is soluble in water so remains in aqueous solution, and no solid will remain in excess dilute sulfuric acid. However, the sulfates of the other Group 2 elements tend to form an insoluble sulfate layer on the carbonate, preventing further reaction after the initial effervescence of carbon dioxide gas is seen.

The reactions with dilute hydrochloric acid also form soluble salts, the chlorides.

Thermal decomposition of Group 2 carbonates and nitrates

The carbonates and nitrates of the Group 2 elements decompose when heated. The carbonates break down to form the metal oxide and give off carbon dioxide gas.

The temperature at which thermal decomposition takes place **increases** going down Group 2. The Group 2 nitrates also undergo thermal decomposition.

Group 2 nitrates emit toxic nitrogen dioxide, NO_2, when heated. Higher temperatures are required for thermal decomposition, like carbonates.

Some uses of Group 2 compounds

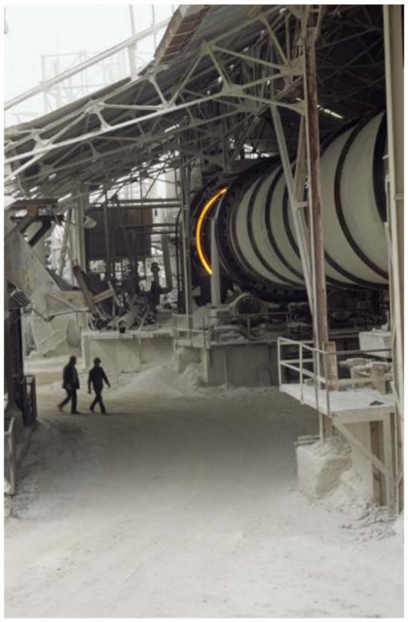


Figure 11.8 In a rotating lime kiln, calcium carbonate undergoes thermal decomposition to form calcium oxide and carbon dioxide.

Limestone, primarily composed of calcium carbonate, is used for building blocks and mortar. It is now commonly used alongside cement and sand, although cement is also made from lime. Marble is another calcium carbonate used for building materials. The first stage of cement production involves heating limestone in a lime kiln to form calcium oxide.

$$CaCO_3(s) \xrightarrow{\text{heat}} \underset{\text{lime}}{CaO(s)} + CO_2(g)$$

Calcium oxide from lime kiln is used to make cement, a widely used building material, and slaked lime for raising soil pH. Cement is mixed with sand and rock for concrete, while slaked lime neutralizes acid.

Chapter 12 – Group 17

Physical properties of Group 17 elements

In this chapter we will look at the elements in Group 17 of the Periodic Table, called the **halogens**.

Here are the electronic configurations of the first four elements in Group 17:

Fluorine (F)	$1s^2 2s^2 2p^5$
Chlorine (Cl)	$1s^2 2s^2 2p^6 3s^2 3p^5$
Bromine (Br)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$
Iodine (I)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$

Group 17 elements are non-metals, forming diatomic molecules with two atoms at room temperature. Each molecule has a single covalent bond, and their physical properties are illustrated in Table 12.

Group 17 element	Atomic radius / nm	Melting point / °C	Boiling point / °C	Colour
fluorine (F_2)	0.072	−220	−188	pale yellow
chlorine (Cl_2)	0.099	−101	−35	green/yellow
bromine (Br_2)	0.114	−7	59	orange/ brown
iodine (I_2)	0.133	114	184	grey/black solid, purple vapour

Table 12.1 Some physical properties of the Group 17 elements. The atomic radius value is taken from single covalent data (see page 51).

Figure 12.2 Three of the Group 17 elements, known as the halogens.

The melting and boiling points of halogens increase as the number of electrons in their molecules increases. This increases the volatility of halogens, as they evaporate more easily. The van der Waals' forces between molecules also increase, making iodine less volatile and fluorine the most volatile. The colors of halogens also darken as the group size increases.

Reactions of Group 17 elements

The halogen atoms need to gain just one more electron to achieve the stable electronic configuration of the noble gas atoms to the right of them in the Periodic Table. Therefore, they react with metallic elements, with each of their atoms gaining an electron from a metal atom to become ions with a 1− charge.

Halogens, when reacting with metal atoms, gain one electron, becoming electron acceptors. They are reduced in the process, reducing their oxidation number from 0 to -1.

Halogens react with numerous non-metals, each atom sharing electrons in a covalent bond, such as in hydrogen chloride (HCl).

Fluorine atoms have stronger electron pull in covalent bonds, compared to iodine atoms. This is due to their smaller size, closer outer shell, and less shielding from positive nuclear charge attraction. Fluorine's 9+ nuclear charge makes it a stronger oxidizing agent.

Halogen	Electronegativity
fluorine (F)	4.0
chlorine (Cl)	3.0
bromine (Br)	2.8
iodine (I)	2.5

Table 12.2 Electronegativity values of the halogens.

Hydrogen–halogen bond	Bond energy / kJ mol^{-1}
H—F	562
H—Cl	431
H—Br	366
H—I	299

Table 12.4 Hydrogen–halogen bond energies.

The halogens get less reactive going down Group 17.

Reactions with hydrogen

The halogens form hydrogen halides with hydrogen gas. One example of this is the reaction of hydrogen with chlorine to give hydrogen chloride (see page 173). The trend in reactivity is illustrated by their reactions (Table 12.3).

Hydrogen halides have varying thermal stability, and hydrogen iodide can be decomposed by inserting a red-hot wire into the gas, resulting in purple fumes.

$$2HI(g) \longrightarrow H_2(g) + I_2(g)$$

Hydrogen fluoride and chloride are stable up to 1500°C, while hydrogen bromide is more resistant to decomposition than hydrogen iodide. At 430°C, 10% of HBr decomposes, while 20% of HI does.

The hydrogen halides get less thermally stable going down Group 17.

most thermally stable HF
HCl
HBr
least thermally stable HI

Equation for reaction	Description of reaction
$H_2(g) + F_2(g) \longrightarrow 2HF(g)$	reacts explosively even in cool, dark conditions
$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$	reacts explosively in sunlight
$H_2(g) + Br_2(g) \longrightarrow 2HBr(g)$	reacts slowly on heating
$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$	forms an equilibrium mixture on heating

Table 12.3 The reactions of hydrogen and the halogens, showing decreasing reactivity going down Group 17.

Reactions of halide ions with concentrated sulfuric acid

We can prepare hydrogen chloride gas by dropping concentrated sulfuric acid slowly onto crystals of sodium chloride (see the apparatus in Figure 12.5):

$$NaCl(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HCl(g)$$

The reaction produces white fumes of HCl but cannot be used to prepare pure hydrogen bromide or hydrogen iodide. Hydrogen halides decompose, oxidizing HBr and HI. Concentrated sulfuric acid, a strong oxidizing agent, can oxidize and decompose these, causing further reactions.

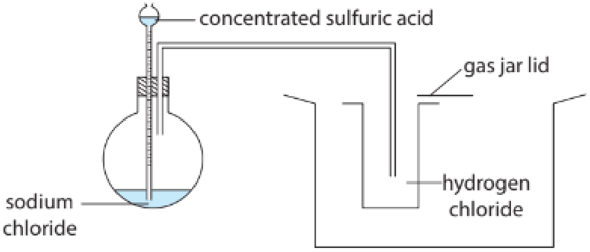


Figure 12.5 Preparing a sample of hydrogen chloride gas. The hydrogen chloride gas is denser than air so displaces the air from the gas jar as it collects.

Several observations can be made here:

- Sulfur is seen as a yellow solid
- Hydrogen sulfide has a strong smell of bad eggs
- Iodine is produced as a violet/purple vapour.

It gets easier to oxidise the hydrogen halides going down Group 17.

$$NaBr(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HBr(g)$$

followed by oxidation of HBr(g):

$$2HBr(g) + H_2SO_4(l) \longrightarrow Br_2(g) + SO_2(g) + 2H_2O(l)$$

A reddish brown gas is seen; this is the element bromine.

With sodium iodide, the sulfuric acid is reduced to a variety of sulfur products as it oxidises the hydrogen iodide formed by different degrees. The products of the oxidation of HI are sulfur dioxide, sulfur and hydrogen sulfide, as shown in the reactions below:

$$NaI(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HI(g)$$

followed by oxidation of HI(g):

$$2HI(g) + H_2SO_4(l) \longrightarrow I_2(g) + SO_2(g) + 2H_2O(l)$$

and:

$$6HI(g) + H_2SO_4(l) \longrightarrow 3I_2(g) + S(s) + 4H_2O(l)$$

and:

$$8HI(g) + H_2SO_4(l) \longrightarrow 4I_2(g) + H_2S(g) + 4H_2O(l)$$

Disproportionation

The element chlorine (Cl2, oxidation number = 0) undergoes a type of redox reaction called **disproportionation** when it reacts with alkali. Disproportionation can be thought of as a ‘self-reduction/ oxidation’ reaction.

Chlorine in cold alkali (15°C)

$$Cl_2(aq) + 2NaOH(aq) \longrightarrow NaCl(aq) + \underset{\text{sodium chlorate(I)}}{NaClO(aq)} + H_2O(l)$$

The ionic equation for the reaction is:

$$\underset{0}{Cl_2}(aq) + 2OH^-(aq) \longrightarrow \underset{-1}{Cl^-}(aq) + \underset{+1}{ClO^-}(aq) + H_2O(l)$$

oxidation number of Cl

The ionic equation for this redox reaction can be split into two half-equations, showing the reduction and oxidation.

- The **reduction** reaction (in which chlorine’s oxidation number is reduced is):

$$\underset{0}{\tfrac{1}{2}Cl_2} + e^- \longrightarrow \underset{-1}{Cl^-}$$

oxidation number of Cl

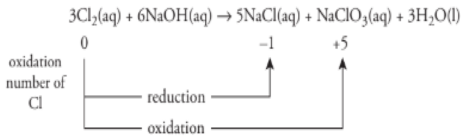
- The **oxidation** reaction is:

$$\underset{0}{\tfrac{1}{2}Cl_2} + 2OH^- \longrightarrow \underset{+1}{ClO^-} + H_2O + e^-$$

oxidation number of Cl

Chlorine in hot alkali (70°C)

When we add chlorine and hot concentrated aqueous sodium hydroxide a different disproportionation reaction takes place:



$$3Cl_2(aq) + 6NaOH(aq) \rightarrow 5NaCl(aq) + NaClO_3(aq) + 3H_2O(l)$$

Uses of the halogens and their compounds

Chlorination of water

Adding a small amount of chlorine to a water supply will kill bacteria and make the water safer to drink. The chlorine undergoes disproportionation in water:

$$\begin{array}{lccccccc} & Cl_2(aq) & + & H_2O(l) & \longrightarrow & HCl(aq) & + & HClO(aq) \\ \text{oxidation number of Cl} & 0 & & & & -1 & & +1 \end{array}$$

HClO is called chloric(I) acid. It decomposes slowly in solution. One theory suggests that it produces reactive oxygen atoms that can kill bacteria in water:

$$HClO \longrightarrow HCl + [O]$$

Bleach

Bleach is an equal mixture of sodium chloride (NaCl) and sodium chlorate(I) (NaClO), made from chlorine and cold alkali. It ‘bleaches’ colours and stains because oxygen atoms from the chlorate(I) ions oxidise dye and other coloured molecules. They also kill bacteria when toilets are cleaned with bleach (see Figure 12.6).

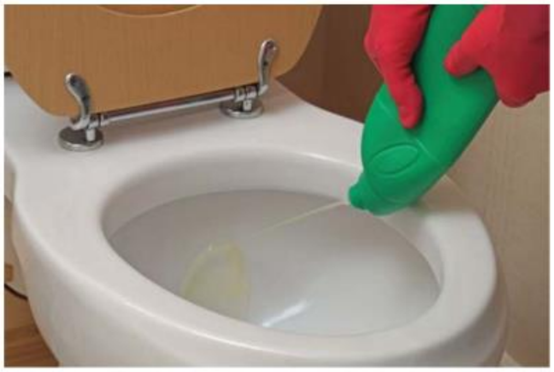


Figure 12.6 The reaction: $Cl_2(aq) + 2NaOH(aq) \longrightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$ is used in industry to produce bleach. The bleaching agent is the chlorate(I) ion.

Other uses

Halogens are present in organic compounds like PVC and halogenated hydrocarbons used as solvents, refrigerants, and aerosols.

Revision questions

1) Magnesium and strontium are in Group 2 of the Periodic Table.

(i) When reacted with oxygen, magnesium forms a white powder called magnesium oxide. Write the equation for the reaction of magnesium with oxygen.

(ii) Magnesium reacts with dilute acids. Describe what you would expect to see when magnesium ribbon is added to an excess of dilute hydrochloric acid.

(iii) Strontium reacts in a similar way to magnesium. Describe one difference you might observe if strontium, instead of magnesium, was reacted with dilute hydrochloric acid

2) The Group 2 element barium, Ba, is silvery white when pure but blackens when exposed to air. The blackening is due to the formation of both barium oxide and barium nitride.

The nitride ion is N^{3-}

(a) Predict the formula of:

barium oxide .... barium nitride ....

(b) A 0.11 g sample of pure barium was added to 100 cm^3 of water.

$$Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$$

(i) Show that 8.0×10^{-4} mol of Ba were added to the water.

(ii) Calculate the volume of hydrogen, in cm^3, produced at room temperature and pressure.

(iii) Calculate the concentration, in mol dm^{-3}, of the $Ba(OH)_2(aq)$ solution formed.

(iv) State the approximate pH of the $Ba(OH)_2(aq)$ solution.

(c) A student repeated the experiment in (b) using a 0.11 g sample of barium that had blackened following exposure to the air. Suggest why the volume of hydrogen produced would be slightly less than the volume collected using pure barium.

(d) Describe and explain the trend, down the group, in the reactivity of the Group 2 elements with water.

3) When heated strongly, $CaCO_3$ decomposes. Write an equation, including state symbols, for the thermal decomposition of $CaCO_3$.

4. Calcium oxide reacts with water and with nitric acid. State the formula of the calcium compound formed when:

(i) calcium oxide reacts with water,
(ii) calcium oxide reacts with nitric acid

5) A student prepared an aqueous solution of calcium chloride by reacting calcium with hydrochloric acid.

$$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$$

(i) Using oxidation numbers, show that this is a redox reaction.
(ii) The student had added the exact amount of calcium required to react with the hydrochloric acid used. After carrying out the experiment, the student accidentally added some more calcium. The student was surprised that the extra calcium still reacted. Explain this observation. Include an equation in your answer.

6) Barium reacts with water in a redox reaction.

$$Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$$

(i) Explain, in terms of electrons, what is meant by oxidation.
(ii) Which element has been oxidised in this reaction? Deduce the change in its oxidation number.

element ....

oxidation number changes from .... to............

7) Magnesium reacts with oxygen to form magnesium oxide.
$2Mg(s) + O2(g) \rightarrow 2MgO(s)$
(i) Use oxidation numbers to show that oxygen has been reduced in its reaction with magnesium.
(ii) Draw a 'dot-and-cross' diagram to show the arrangement of electrons in magnesium oxide. Show outer electron shells only and include any charges.

8) Old samples of magnesium oxide become contaminated with magnesium carbonate. (i) Suggest how this contamination takes place.

(ii) A student added an excess of hydrochloric acid to an old sample of magnesium oxide that is contaminated with magnesium carbonate. State two observations that the student would make.

(iii) Explain, with the aid of equations, why the resulting solution contained only one
dissolved compound of magnesium.

9) **Iodine is used in many inorganic and organic reactions.**

(a) (i) State and explain the trend in volatility of the halogens, from chlorine to iodine.[2]

(ii) Explain why HI is the least thermally stable of HCl, HBr and HI.[1]

(iii) The table shows the electronegativity values for hydrogen, fluorine and iodine.

element	electronegativity value
H	2.1
F	4.0
I	2.5

Explain, in terms of intermolecular forces, why HI has a lower boiling point than HF.[2]

(iv) Iodine reacts with hot concentrated aqueous sodium hydroxide in the same way as chlorine.

Write an equation for the reaction of iodine and hot aqueous sodium hydroxide. [1]

(b) Iodoalkanes contain carbon-iodine bonds.

The simplest iodoalkane is CH_3I.

(i) CH_3I can be made from methanol, CH_3OH.

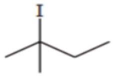
Identify a reagent that can convert CH_3OH to CH_3 I.[1]

(ii) 1,2-diiodoethane, CH_2ICH_2I, can be made by bubbling ethene into liquid iodine.

Fully name the type of mechanism shown in this reaction.[1]

(c) J reacts with NaOH, forming different products dependent on the conditions used.

J

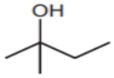


(i) Name J.

.... [1]

(ii) J reacts with NaOH(aq) to form K.

K



Fully name the mechanism of the reaction of J with NaOH(aq) to form K.

.... [1]

10)

The Group 17 elements, chlorine, bromine and iodine, are non-metals that show trends in their physical and chemical properties.

(a) Describe the trend in the colour of the Group 17 elements down the group.[1]

(b) The Group 17 elements can oxidise many metals to form halides.
(i) Describe the relative reactivity of the elements in Group 17 as oxidising agents.[1]
(ii) Chlorine reacts with hot tin metal to form tin(IV) chloride, $SnCl_4$.
$SnCl_4$ is a colourless liquid at room temperature that reacts vigorously with water to form an acidic solution.
Suggest the type of structure and bonding shown by $SnCl_4$. Explain your answer.[2]

(c) The Group 17 elements form soluble halides with sodium.
(i) Describe what is seen when dilute $AgNO_3(aq)$ is added to $NaBr(aq)$ followed by aqueous ammonia. [2]

(ii) $NaCl$ reacts with concentrated H_2SO_4 to form HCl and $NaHSO_4$.
Explain the difference between the reactions of concentrated H_2SO_4 with $NaCl$ and with NaI. Your answer should refer to the role of the sulfuric acid in each reaction.[3]

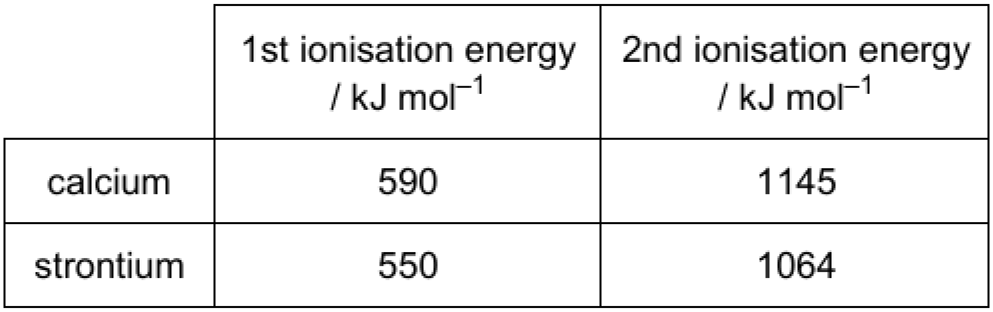
(d) The hydrogen halides are useful reagents in organic and inorganic reactions.
(i) Describe and explain the trend in the boiling points of the hydrogen halides, HCl, HBr and HI.[2]
(ii) Describe and explain the trend in the thermal stabilities of the hydrogen halides, HCl, HBr and HI.

(e) Lucas's reagent is a mixture of HCl and $ZnCl_2$. Primary, secondary and tertiary alcohols can be distinguished by their reaction with Lucas's reagent.
Alcohols react with the HCl in Lucas's reagent to form halogenoalkanes.
$ZnCl_2$ acts as a homogeneous catalyst for these reactions.
(i) Explain the meaning of the term homogeneous. [1]
(ii) Pentan-3-ol, $C_2H_5CH(OH)C_2H_5$, reacts slowly with HCl to form a secondary halogenoalkane.
Complete the equation for this reaction using structural formulae.
$C_2H_5CH(OH)C_2H_5+$

(iii) The fastest reaction shown by Lucas's reagent is with a tertiary alcohol.
Draw the structure of the tertiary alcohol that is an isomer of pentan-3-ol.[1]

(iv) Tertiary alcohols tend to react with Lucas's reagent using the same mechanism as in their reaction with HCl.
Suggest the type of reaction shown by tertiary alcohols with Lucas's reagent.[1][Total: 17]

11) In their reactions, calcium and strontium each lose electrons to form ions with a 2+ charge. The first and second ionisation energies of calcium and strontium are shown below.



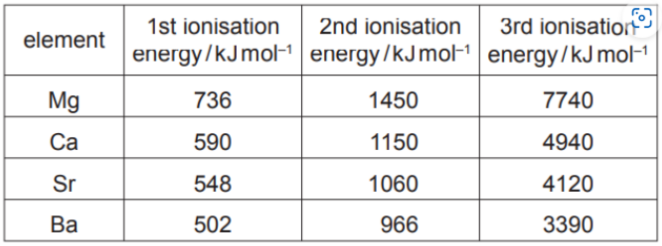
	1st ionisation energy / kJ mol^{-1}	2nd ionisation energy / kJ mol^{-1}
calcium	590	1145
strontium	550	1064

(i) Write an equation, with state symbols, to represent the second ionisation energy of calcium.
(ii) Why are the second ionisation energies of calcium and strontium greater than their first ionisation energies?
(iii) Explain why the first and second ionisation energies of strontium are less than those of calcium.

12) Carbon dioxide can be prepared easily in the laboratory by the action of heat on most carbonates. Construct an equation to illustrate this reaction

13)

(a) Successive ionisation energies for the elements magnesium to barium are given in the table



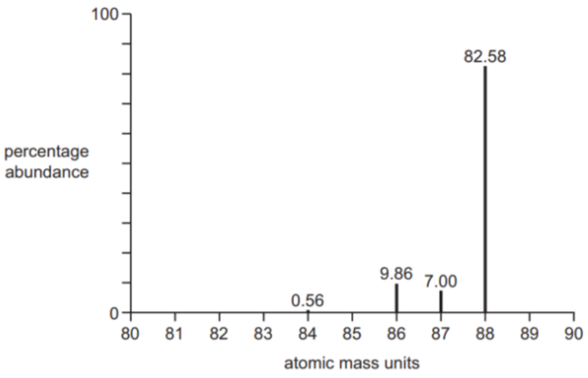
element	1st ionisation energy / kJ mol^{-1}	2nd ionisation energy / kJ mol^{-1}	3rd ionisation energy / kJ mol^{-1}
Mg	736	1450	7740
Ca	590	1150	4940
Sr	548	1060	4120
Ba	502	966	3390

(i) Explain why the first ionisation energies decrease down the group.

.... [3]

(ii) Explain why, for each element, there is a large increase between the 2nd and 3rd ionisation energies.

(b) A sample of strontium, atomic number 38, gave the mass spectrum shown. The percentage abundances are given above each peak.



(i) Complete the full electronic configuration of strontium.

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

(ii) Explain why there are four different peaks in the mass spectrum of strontium. [1]

(iii) Calculate the atomic mass, A_r, of this sample of strontium.

Give your answer to three significant figures.

$A_r =$

(c) A compound of barium, **A**, is used in fireworks as an oxidising agent and to produce a green colour.

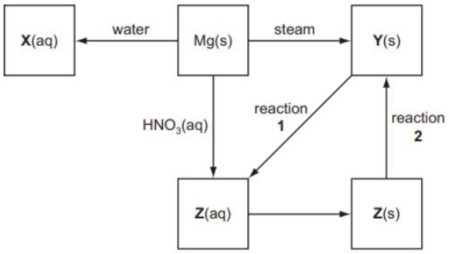
(i) Explain, in terms of electron transfer, what is meant by the term oxidising agent. [1]

(ii) A has the following percentage composition by mass: Ba, 45.1; Cl, 23.4; O, 31.5.

Calculate the empirical formula of **A**.

empirical formula of **A** [3]

(d) Some reactions involving magnesium and its compounds are shown in the reaction scheme below.



(i) Give the formulae of the compounds **X**, **Y** and **Z**.

x

Y

Z [3]

(ii) Name the reagent needed to convert Y(s) into Z(Zq) in reaction 1 and write an equation for the reaction.

reagent

equation [2]

(iii) How would you convert a sample of Z(s) into Y(s) in reaction 2 ?

[1]

Mg to **X**

Z to Y

14)

The elements in Group 2, and their compounds, show many similarities and trends in their properties.

(a) Magnesium, calcium, strontium and barium all react with cold water.

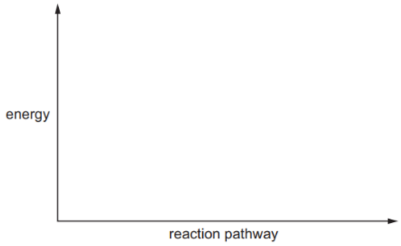
(i) Describe what you would see when some calcium is added to cold water.[3]

(ii) Write an equation for the reaction taking place in (i).[1]

(iii) Describe how the reaction of barium with cold water would differ from the reaction of calcium in (i) in terms of what you would see.[1]

(b) Magnesium oxide can be formed by the reaction of magnesium and oxygen in the air.

(i) Draw a fully labelled reaction pathway diagram for the reaction between magnesium and oxygen.



(ii) Explain why there is no visible reaction when a piece of magnesium ribbon is exposed to the air

(ii) Explain why there is no visible reaction when a piece of magnesium ribbon is exposed to the air

(iii) Magnesium oxide is used to manufacture heat-resistant bricks for furnace linings in the steel-making industry. State and explain the property of magnesium oxide that makes it suitable for this use.[2]

(iv) Suggest a reason why magnesium oxide cannot be used as a lining for any furnaces containing acidic materials.[1]

(c) The nitrates and carbonates of the Group 2 elements, from magnesium to barium, decompose when heated.

(i) State the trend in the temperature of thermal decomposition of these Group 2 nitrates and carbonates.[1]

(ii) Give the equation for the thermal decomposition of magnesium carbonate.[1]

(iii) Give the equation for the thermal decomposition of calcium nitrate.[1][Total: 15]