CHEMICAL FAMILIES; PATTERNS IN PROPERTIES

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Organizer 📌

Objectives

By the end of this topic, the learner should be able to:

(a) Identify and write electron arrangement of alkali metals, alkaline earth metals, halogens and noble gases.
(b) State and explain the trends in physical properties of elements in group I, II, VII and VIII.
(c) State and explain the trends in reactivity of elements in group I, II, VII and VIII.
(d) Explain the similarities in chemical formulae of compounds of the elements in a group.
(e) Explain the unreactive nature of group VIII elements.
(f) Identify and write electron arrangement of period 3 elements.
(g) State and explain the trends in physical and chemical properties of the elements in period 3.

CHEMICAL FAMILIES; PATTERNS IN PROPERTIES

Elements in the same group are said to belong to the same chemical family.

Trends in physical and chemical properties provide useful information in predicting the physical and chemical behaviour of the elements within a family.

1. The Alkali Metals

The elements in **group I** of the periodic table are called **Alkali metals.**

These include, lithium, sodium, potassium, rubidium, caesium and francium. The electron arrangements of the first three alkali metals are as follows:

Lithium (L) : 2.1

Sodium (Na) : 2.8.1

Potassium (k) : 2.8.8.1

Each alkali metal atom has one electron in the outermost energy level. Down the group there is an increase in the number of occupied energy levels.

Task: Draw the atomic structure of the first 3 Alkali Metals.

Gradation in Size of the Atom and Ion

It is not possible to measure the sizes of atoms and ions of elements in the laboratory due to their small size. The table below gives a summary of the atomic and ionic radii of the Alkali Metals.

Element	Symbol	Atomic number	Atomic radius (nm)	Ionic radius (nm)
Lithium	Li	3	0.133	0.060
Sodium	Na	11	0.157	0.095
Potassium	K	19	0.203	0.133

Atomic radius *is the distance between the centre of the nucleus of an atom and the outermost energy level occupied by an electron or electrons*.

Discussion Questions

1. **State and explain the trends in the atomic and ionic radii down the group**

The atomic radii and ionic radii of the alkali metals increase down the group. This is because each alkali metal has one more occupied energy level than the preceding member in the group.

Lithium has two energy levels. Sodium has three while potassium has four. The outermost electron in a sodium atom is therefore further from the nucleus than the outermost electron in a lithium atom. This explains the increase in the atomic and ionic radii down the group.

1. **How does the ionic radius and atomic radius of an element compare?**

The ionic radius of an alkali metal is **less than its atomic radius.**

An alkali metal forms an ion by losing the single electron from the outermost energy level. The resulting ion (cation) has one occupied energy level less than the corresponding atom.

When an atom loses an electron to form a positively charged ion, the remaining electrons experience greater nuclear attraction. The remaining energy levels move closer to the nucleus resulting in a reduction in the radius.

Physical Properties of Alkali metals

The table below shows some physical properties of alkali metals.

Element	Appearance	Ease of cutting	Melting point (°C)	Boiling point (°C)	Electrical conductivity	Atomic radius (nm)	1st Ionisation energy ($kJmol^{-1}$)
Lithium	Slivery white	Slightly hard	180	1330	Good	0.133	520
Sodium	Shiny grey	Easy	98	890	Good	0.157	496
Potassium	Shiny grey	Easy	64	774	Good	0.203	419

Ionization energy *is the minimum energy required to remove an electron from the outermost energy level of an atom in the gaseous state*.

Discussion Questions

State and explain the trends in the following properties down the group?

(i) Appearance

The alkali metals have a **shiny metallic lustre** when freshly cut. However, the surface quickly **tarnishes**. The surface tarnishes because of **reacting with air**.

(ii) Ease of cutting

The alkali metals are **soft and easy to cut.**

The softness and ease to cut **increases down the group** due to the **decrease** in the strength of the **forces holding the atoms** together as you move down the group.

(iii) Melting and boiling points.

The alkali metals have relatively **low melting and boiling** points.

The melting and boiling points **decrease down the group** due to the **weakening** of the forces holding the atoms together.

The strength of the forces holding atoms together depends on the **size of the atoms**. The **larger** the atoms, the **weaker** the force. Thus **as the atomic radius increases the forces of attraction between the atoms weaken**, hence the decrease in the melting and boiling points down the group.

(iv) Electrical conductivity.

Alkali metals are **good conductors of heat and electricity.** Conductivity in metals is due to the **presence of delocalised electrons** in the structure of the metal. Since they all have one electron in their outermost energy level, their conductivity is **similar**.

In metals, the electrons in the outermost energy level move randomly throughout the metallic structure. Since the electrons do **not remain in one fixed position**, they are said to be **delocalised**.

(v) 1st ionization energy

Down the group from lithium to potassium, the **1st ionization energy decreases.** This means that **less** energy is needed to remove the electron from the outermost energy level of a **potassium atom than a sodium atom** and a **lithium** atom. This is because the **effective force of attraction on the outermost electron** by the positive nucleus **decreases with increasing atomic size** and **distance from the nucleus.**

Chemical Properties of Alkali Metals

Alkali metals react by **losing the one electron** from their **outermost** energy level to attain a **stable electron configuration**. Their reactivity **increases down the group**.

Reaction with air and water

- Lithium, sodium and potassium **react vigorously** with both air and water. Potassium is the **most reactive** and lithium, the **least reactive.**
- When sodium is exposed, it reacts with moisture in the air to form sodium hydroxide. The sodium hydroxide further reacts with carbon (IV) oxide in the air to form sodium carbonate.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

$$2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3 \cdot H_2O(s)$$

- Sodium **burns in air with a yellow flame** to form a yellowish white solid which is mainly sodium oxide.

$$\text{Sodium + Oxygen} \longrightarrow \text{Sodium Oxide}$$

$$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

- When sodium burns in **air enriched with oxygen** it forms mainly **sodium peroxide.**

$$\text{Sodium + Oxygen} \longrightarrow \text{Sodium peroxide}$$

$$2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$$

- Potassium **burns in air with a lilac flame** to form a white solid which is potassium oxide.

$$\text{Potassium + oxygen} \longrightarrow \text{Potassium oxide}$$

$$4K(s) + O_2(g) \longrightarrow 2K_2O(s)$$

The observations made when alkali metals react with water are summarised in the table below.

Metal	Observation when metal reacts with water	Rate of reaction
Lithium	Lithium floats in water. A colourless gas is produced. The gas does not ignite spontaneously. The resulting solution turns red litmus paper blue.	Vigorous
Sodium	Sodium metal darts on the water surface as it melts into a silvery ball. A hissing sound is produced. A colourless gas is produced which may ignite spontaneously. The solution formed is alkaline.	More vigorous

Potassium	The metal darts about on the surface of the water and melts into a silvery ball. A colourless gas is produced which spontaneously bursts into a flame. Potassium vapour burns with a lilac flame. The resulting solution is alkaline.	Explosive

- **Alkali metals react with water to form alkaline solutions and hydrogen gas.**

Metal + Water ⟶ Metal hydroxide + Hydrogen

Lithium + Water ⟶ Lithium hydroxide + hydrogen gas

$$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$$

Sodium + Water ⟶ Sodium hydroxide + Hydrogen gas

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Potassium + water ⟶ Potassium hydroxide + Hydrogen gas

$$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$$

- Potassium is **the most reactive alkali metal**. This is because the **electron** in the outermost energy level is **loosely held by the nucleus** and is **easily removed during a reaction.**
- The ease of losing valence electrons **increases** down the group as the **atomic radius increases,** hence the **increase** in reactivity from lithium to potassium.

Reaction with chlorine

Alkali metals react with chlorine gas to form the corresponding **metal chlorides.** The reactivity of alkali metals with chlorine ***increases down the group.*** This is because of the **increase in atomic radius** which leads to **increasing ease to lose the electron** in the outermost energy level.

Task: Cut a small piece of sodium and place it in a deflagrating spoon. Warm it and quickly lower it into a gas jar containing chlorine. Record your observations.

Discussion Questions

1. **What is observed when a hot piece of sodium metal is lowered into a gas jar containing chlorine?**

   When hot sodium metal is lowered into chlorine gas, the metal **bursts into flame,** white fumes of sodium chloride are formed.

2. **What is formed when sodium metal reacts with chlorine?**

   Sodium metal + chlorine gas ⟶ Sodium chloride

   $$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

3. **Predict how lithium and potassium would react with chlorine.**

   Lithium reacts **less vigorously** with chlorine while potassium **reacts much more violently** with chlorine than sodium.

   Lithium + chlorine gas ⟶ Lithium chloride

   $$2Li(s) + Cl_2(g) \longrightarrow 2LiCl(s)$$

   Potassium + chlorine gas ⟶ Potassium chloride

   $$2K(s) + Cl_2(g) \longrightarrow 2KCl\ (s)$$

Similarity of Ions and formulae of some compounds of Alkali metals.

Formulae of hydroxides, oxides and Chlorides of Alkali metals

Alkali metal ion	Hydroxide (OH^-)	Oxide (O^{2-})	Chloride (Cl^-)
Li^+			LiCl
Na^+		Na_2O	
K^+	KOH		

Discussion Question

Explain the similarity in the chemical formulae of the compounds formed between alkali metal ions and hydroxides, oxides and chloride ions.

Each alkali metal ion combines with a **single hydroxide ion** to form the respective hydroxide namely; Lithium hydroxide (LiOH), sodium hydroxide (NaOH) and potassium hydroxide (KOH). **This is because the valency of Group I elements is one**. The same applies in the formation of lithium chloride (LiCl), sodium chloride (NaCl) and potassium chloride (KCl)

Two alkali metal ions combine with **one oxide ion** to form the corresponding oxide namely; Lithium oxide (L_2O), sodium oxide (Na_2O) and potassium oxide (K_2O).

This is because the valency of **oxygen is two.** Therefore **one oxygen ion requires two alkali metal ions** to combine with to form the corresponding oxide.

Due to their high reactivity, alkali metals are not found as free elements. They are normally found in the combined state in the earths' crust.

Uses of Alkali metals and their compounds

1. Sodium is used in the manufacture of Sodium cyanide for use in the extraction of gold.
2. Lithium is used in the manufacture of special high strength glasses and ceramics
3. Lithium compounds are used in the manufacture of dry cells for use in mobile phones, laptops, stop watches and zero emission electric vehicles.
4. A molten mixture of sodium and potassium is used as a coolant in nuclear reactors.
5. Sodium vapour is used to produce the yellow glow in streetlights.
6. Molten sodium is used as a reducing agent in the extraction of titanium from titanium (IV) chloride.

$$TiCl_4(g) + 4Na(l) \longrightarrow Ti(s) + 4NaCl(l)$$

7. Sodium chloride is used as a food additive.
8. A mixture of Sodium hydroxide (caustic soda) and carbon disulphide is used in the manufacture of artificial silk called rayon.

2. Alkaline Earth Metals

The elements in **group II** of the periodic table are called **alkaline earth metals.**

They consist of beryllium, magnesium, calcium, strontium, barium and radium. The electron arrangement of the first three alkaline earth metals is as follows:

Beryllium (Be) ; 2.2

Magnesium (Mg); 2.8.2

Calcium (Ca) ; 2.8.8.2

An atom of an alkaline earth metal has two electrons in the outermost energy level.

Task: Draw the atomic structure of the first 3 Alkaline Earth Metals

Gradation in Size of Atom and Ion

The table below summaries the atomic and ionic sizes of the Alkaline Earth Metals.

Element	Symbol	Atomic number	Electron arrangement	Atomic radius (nm)	Ionic radius (nm)
Beryllium	Be	4	2.2	0.089	0.031
Magnesium	Mg	12	2.8.2	0.136	0.065
Calcium	Ca	20	2.8.8.2	0.174	0.099

Discussion Question

State and explain the trends in atomic and ionic sizes down the group

Among the alkaline earth metals, the **atomic radius increases down** the group as **more energy levels are occupied.**

Beryllium has the **smallest** atomic radius among the alkaline earth metals because it has the **least number of occupied energy levels.**

Group II elements form ions by losing the two electrons in the outermost energy level in order to attain a stable electron arrangement. The loss of two electrons in the outermost energy level accounts for the smaller ionic radius compared to the atomic radius of the corresponding atom.

Ion	Electron arrangement
Be^{2+}	2
Mg^{2+}	2.8
Ca^{2+}	2.8.8

Beryllium ion with only one occupied energy level is therefore the smallest ion.

Physical Properties of the Alkaline Earth Metals

Element	Atomic number	Melting point (°C)	Boiling point (°C)	Atomic radius (nm)	1st I.E ($kJmol^{-1}$)	2nd I. E ($kJmol^{-1}$)
Beryllium	4	1280	2450	0.089	900	1800
Magnesium	12	650	1110	0.136	736	1450
Calcium	20	850	1140	0.174	590	1150
Strontium	38	789	1330	0.210	550	1060
Barium	56	725	1140	0.220	503	970

The table below summarises the Physical Properties of Alkaline Earth Metals

Discussion Questions

1. **What is the appearance of a polished surface of an alkaline earth metal?**

- When an alkaline earth metal is polished, it acquires a metallic lustre.

- Alkaline earth metals lose their metallic lustre when **exposed to air** because of **oxidation**.
- The purpose of **polishing the surface** of the alkaline earth metals before using them in experiments is to **remove the thin oxide layer** that usually forms on their surface. Both **magnesium and calcium have a metallic lustre.**

2. **What is observed when:**

   (a) One cuts an alkaline earth metal?

- Magnesium is hard to cut with a knife. However it is ductile and malleable. Calcium is brittle hence it cannot be cut with a knife.

  A **ductile** material is one which can be drawn into a wire. Materials which **can be hammered into sheets** are said to be **malleable**. A **brittle** substance is one which is **hard and likely to break.**

i. **Magnesium and calcium are separately used to complete an electric circuit?**

Both magnesium and calcium are **good conductors of heat and electricity** due to the presence of **delocalised electrons.**

3. **How do the physical properties of alkaline earth metals compare with those of alkali metals?**
4. **State and explain the trend in:**

   (a) Melting and boiling points

The melting and boiling points of beryllium are **very high** compared to other alkaline earth metals. This is because the **beryllium atom is very small** and the forces of attraction between the atoms are **very strong.**

Down the group the melting point and boiling points **decrease**. This is because in metals atoms are held together by **forces of attraction between positive nuclei and delocalised electrons.** As the atomic radius increase this **attraction decreases** because of the **increasing distance from the positive nucleus** to the delocalised electrons. This explains why the melting point and boiling point decreases down the group.

   (b) Ionization energy down the group.

The first and second ionization energies decreases down the group. This is because the **effective force of attraction on the outermost electron** by the positive nucleus **decreases with increasing atomic size** and **distance from the nucleus.**

The first ionization energy for magnesium is the minimum amount of energy required to remove one electron from the outer most energy level.

$$Mg(g) \longrightarrow Mg^{+}(g) + e^{-} \ (1^{st} \text{ I.E.} = 736 \text{ kJmol}^{-1})$$

The **second ionisation energy** of magnesium is the minimum amount of energy required to remove a **second electron** from a magnesium ion with a single positive charge.

$$Mg^{2+}(g) \longrightarrow Mg^{2+}(g) + e^{-} \ (2^{nd} \text{ I.E.} = 1450 \text{ kJmol}^{-})$$

5. **Explain why the second ionization energy is always higher than the first ionization energy.**

Once an electron has been lost from an atom, the overall positive charge holds the remaining electrons more firmly. This then means that removing a second electron form the ion requires more energy than the first electron.

Chemical Properties of Alkaline Earth Metals

Reaction with air and water

Discussion Questions

1. **What is observed when magnesium and calcium are heated in air?**

Magnesium burns in air with a **blinding brilliant white flame** forming a **white solid**. The white solid is a **mixture of magnesium oxide and magnesium nitride.**

$$\text{Magnesium} + \text{Oxygen} \longrightarrow \text{Magnesium oxide}$$

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

$$\text{Magnesium} + \text{nitrogen} \longrightarrow \text{Magnesium nitride}$$

$$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$$

Calcium on the other hand burns with a **faint orange-red flame** forming a **white solid,** which is a **mixture of calcium oxide and calcium nitride.**

$$\text{Calcium} + \text{Oxygen} \longrightarrow \text{Calcium oxide}$$

$$2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$$

$$\text{Calcium} + \text{Nitrogen} \longrightarrow \text{Calcium nitride}$$

$$3Ca(s) + N_2(g) \longrightarrow Ca_3N_2(s)$$

The trend in reactivity of the alkaline earth metals when burning in air is **not clear** due to the **oxide coating on calcium.**

2. **What is observed when magnesium and calcium are placed in water and solution tested with litmus paper?**

Magnesium **reacts slowly with cold water** to form magnesium hydroxide and hydrogen gas bubbles of which stick on the surface of the metal. Magnesium hydroxide **dissolves slightly** in water to form an **alkaline** solution.

$$\text{Magnesium} + \text{Water} \longrightarrow \text{Magnesium hydroxide} + \text{Hydrogen gas}$$

$$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

A steady stream of hydrogen gas is evolved when calcium reacts with cold water. **A white suspension** appears in the beaker due to the formation of calcium hydroxide which is **sparingly soluble** in water. The calcium hydroxide solution formed is **alkaline**.

$$\text{Calcium} + \text{Water} \longrightarrow \text{Calcium hydroxide} + \text{Hydrogen gas}$$

$$Ca(s) + 2H_2O\ (s) \longrightarrow Ca(OH)_2(aq) + H_2(g)$$

2. **Which of the two metals is more reactive?**

The atomic radii increase from beryllium to calcium. Therefore, the two outer electrons in a calcium atom are more loosely held by the positive nucleus than the outer electrons in magnesium. This means that less energy is required to remove the outer electrons in calcium than in magnesium. Calcium is therefore more reactive than magnesium.

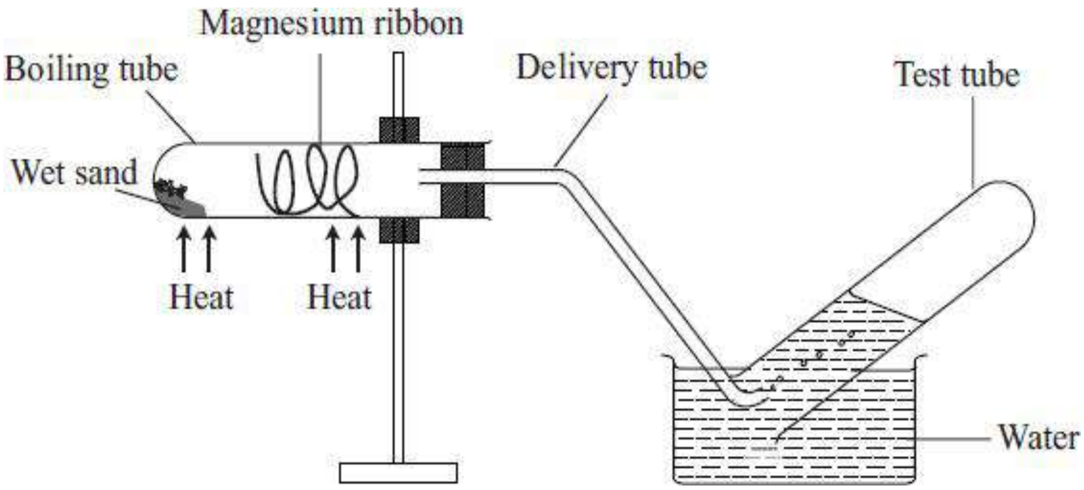
Beryllium
Magnesium ↓ increasing reactivity
Calcium

The order of reactivity of the alkaline earth metals increases down the group.

Reaction with steam

Magnesium reacts slowly with cold water. However, it reacts faster with steam.

To react magnesium with steam, Put some wet sand at the bottom of a test tube. Insert a clean piece of magnesium ribbon (5 cm) into the middle of the test-tube ensure that the coil touches the sides of the tube. Set up the apparatus as shown in below.



Heat the sand gently first, then heat the magnesium ribbon strongly until it glows, then continue to heat the wet sand to generate steam. Record your observations. Test any gas produced using a burning splint. Before the heating is discontinued, the delivery tube should be removed from the water.

Discussion Questions

1. **State and explain what is observed in the reaction tube.**

Magnesium burns in steam with a **white flame.** A **white solid** is formed and a **colourless gas** is produced. A pop sound is produced when a burning splint is introduced into a test-tube containing the gas. Therefore the colourless gas is hydrogen.

$$\text{Magnesium + Steam} \longrightarrow \text{Magnesium oxide + Hydrogen}$$

$$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$$

2. **Why is the wet sand heated initially?**

The sand is heated initially to drive out the air that would otherwise react with magnesium by generating some steam.

3. **Why should the delivery tube be removed before heating is discontinued?**

The delivery tube is removed from the water before heating is stopped at the end of the experiment to prevent sucking back as the apparatus cools.

4. **Explain the observation made when a burning splint is lowered into a test tube of the gas collected.**

A pop sound is produced when a burning splint is introduced into a test-tube containing the gas. Therefore the colourless gas is hydrogen.

Reaction with chlorine

Discussion Questions

1. **State and explain what is observed when chlorine gas reacts with:**
   (a) **A burning piece of magnesium**

When a burning piece of magnesium is lowered into a gas jar containing chlorine, the metal **continues to burn with a brilliant white flame to form a white powder.** The white ash formed is magnesium chloride.

$$\text{Magnesium + Chlorine} \longrightarrow \text{Magnesium chloride}$$

$$Mg(s) + Cl_2(s) \longrightarrow MgCl_2(s)$$

   (b) **A hot piece of calcium**

Calcium may not react steadily with chlorine. This is because a coating of **calcium oxide is formed first when the metal is heated.** However, under suitable conditions calcium may react with chlorine to form calcium chloride.

$$\text{Calcium + Chlorine gas} \longrightarrow \text{Calcium chloride}$$

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

Reaction with dilute acids

Discussion Questions

1. **State and explain what is observed when magnesium ribbon reacts with dilute hydrochloric and sulphuric acids.**

There is **effervescence** when a piece of magnesium is placed in hydrochloric acid. The gas produced during the reaction **produces a 'pop' sound** when a burning splint is introduced to the mouth of the test tube. This shows that the gas produced is **hydrogen**.

$$\text{Magnesium + Hydrochloric acid} \longrightarrow \text{Magnesium chloride + Hydrogen gas}$$

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

$$\text{Magnesium + Sulphuric acid} \longrightarrow \text{Magnesium sulphate + hydrogen gas}$$

$$Mg(s) + H_2SO4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$$

2. **State and explain what is observed when calcium reacts with dilute hydrochloric and sulphuric acids.**

Calcium reacts with dilute hyrdrochloric acid to produce hydrogen gas and calcium chloride.

$$\text{Calcium + Hydrochloric acid} \longrightarrow \text{calcium chloride + hydrogen gas}$$

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

When sulphuric (VI) acid is used, the **reaction quickly stops.** This is due to to the formation of **insoluble calcium sulphate** which **forms a coating on the surface of calcium metal preventing further reaction.**

$$\text{Calcium + Sulphuric acid} \longrightarrow \text{Calcium sulphate + Hydrogen gas}$$

$$Ca(s) + H_2SO_4\,(aq) \longrightarrow CaSO_4(s) + H_2(g)$$

Similarity of ions and formulae of some compounds of alkaline earth metals

All alkaline earth metals have a valency of two (2). Hence the chemical formulae of their compounds are similar.

***Formation of hydroxides, oxides and chlorides of alkaline earth metals*.**

Ion	Hydroxide (OH^-)	Oxide (O^{2-})	Chloride (Cl^-)
Mg		MgO	
Ca	$Ca(OH)_2$		$CaCl_2$

Both magnesium and calcium ions have a **valency of 2.** The hydroxide ion and chloride ion have a **valency of 1.** Therefore **two hydroxide ions combine with one magnesium ion to form magnesium hydroxide, $Mg(OH)_2$.** Similarly, **two chloride ions will combine with one magnesium ion to form magnesium chloride, $MgCl_2$.**

On the other hand, the **oxide ion has valency of 2.** therefore, **one calcium ion combines with one oxide ion to form calcium oxide, CaO.**

(e) Uses of some Alkaline Earth Metals and their Compounds

1. Magnesium is used in the manufacture of magnesium hydroxide which is used as anti-acid medicine. This is because magnesium hydroxide is a non-toxic base.
2. A low-density alloy of magnesium and aluminium is used in aeroplane construction.
3. Hydrated calcium sulphate (Plaster of Paris) is used in hospitals to set fractured bones.
4. Cement is made by heating a mixture of calcium carbonate (limestone), clay and sand.
5. Calcium carbonate is used in the extraction of iron.
6. Calcium oxide (quicklime) is added to acid soils to raise pH for agriculture purposes.
7. Calcium nitrate is used as a nitrogenous fertilizer.
8. Magnesium oxide is used in the lining of furnances.
9. Barium sulphate is used in diagnosis of ulcers.
10. Barium nitrate is used to produce the green flame in fireworks.
11. Calcium carbonate is mixed with oil to make putty.

3. Halogens

The word halogen is derived from the Greek word '*halo'* meaning salt and '*gen'* meaning producer. Halogen thus means salt producer. **Halogens** are non-**metals in Group VII of the periodic table.** Fluorine, chlorine, bromine and iodine are the first four members of the halogen group. The electron arrangement of fluorine and chlorine is as follows:

Fluorine 2.7

Chlorine 2.8.7

Task: Draw the atomic structure of the first 2 Halogens

Gradation in Size of Atoms and Ion in halogens

The table below gives the atomic and ionic sizes of the Halogens.

Halogen	Symbol	Atomic number	Atomic radius (nm)	Ionic radius (nm)
Fluorine	F	9	0.064	0.136
Chlorine	Cl	17	0.099	0.181
Bromine	Br	35	0.114	0.195
Iodine	I	53	0.133	0.216

Discussion Questions

1. **Explain the trends in the atomic and ionic radii of Halogens down the group:**

The atomic and ionic radii of the halogens **increase down the group.** This is because of the **increase** in the number of occupied energy levels.

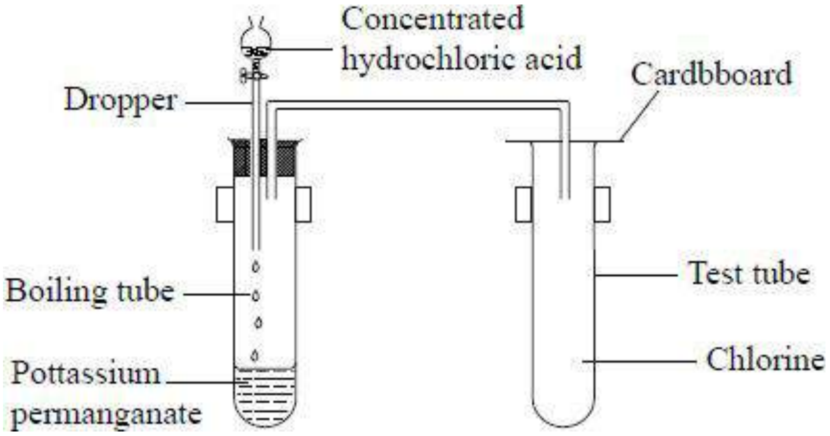
2. **How do the atomic radii of the halogens compare with their ionic radii?**

The **atomic radius of a halogen atom is less than the radius of its ion.**

For example, the **atomic radius of a chlorine atom is less than the ionic radius of a chloride ion.** This is because the **chlorine atom has 17 protons** in the nucleus **attracting 17 electrons** in the energy levels. The chloride **ion has 17 protons in the nucleus attracting 18 electrons.** The effect of the positive nucleus is thus **lower**. **The reduction in the nuclear attraction is due to repulsive effect between the existing electrons and the incoming electron.**

Physical Properties of Halogens

Chlorine can be prepared by the action of concentrated hydrochloric acid on potassium manganate (VII) ($KMnO_4$) or Manganese (IV) oxide (MnO_2). Heat is required when manganese (IV) oxide is used.



Set up for preparing chlorine gas

The table below gives a summary of the physical properties of halogens.

Halogen	Formula	Atomic number	Appearance	Melting point (°C)	Boiling point (°C)
Fluorine	F	9	Pale yellow gas	–238	–188
Chlorine	Cl	17	Greenish yellow gas	–101	–35
Bromine	Br	35	Brown liquid	–7	59
Iodine	I	53	Shiny dark solid	114	184

Discussion Questions

1. **What is the physical state of chlorine, bromine and iodine at room temperature?**

Fluorine and **chlorine** are **gases at room temperature. Bromine** is a **volatile liquid** while **iodine is a solid.**

2. **What is the colour of chlorine, bromine and iodine?**

Fluorine is **pale yellow** while chlorine is **green-yellow.** Bromine is a **brown liquid** while iodine is a **shiny dark grey solid.**

3. **What is observed when chlorine and bromine are shaken with Water and Tetra chloromethane**

When a boiling tube containing chlorine gas or bromine vapour is inverted in water or tetrachloromethane, **the level of the solution rises in the boiling tube.**

The level of the solution rises more in tetrachloromethane than in water. The rise in water level is higher in the case of chlorine compared to bromine. This shows that **chlorine is more soluble in water than bromine. Solubility of halogens in water therefore decreases down the group. All halogens are soluble in tetrachloromethane.**

4. **Explain what is observed when iodine is heated.**

Iodine **sublimes when heated** to form a **purple vapour**. This is because the **particles are held by weak forces** which require **little energy** to break.

5. **Does iodine conduct electricity? Explain**.

Halogens are non conductors of heat and electricity. This is because there are no delocalised electrons in their structures.

6. **Explain the trends in the melting and boiling points of halogens down the group.**

The melting and boiling points of halogens increase down the group.

Halogens exists as **diatomic molecules**. The forces of attraction between molecules (intermolecular forces) **increase with the increase in the size of the molecules.** Hence, the forces of attraction between molecules among the four halogens are **strongest in iodine** and **weakest in fluorine.**

Chemical Properties of Halogens

It is **not easy for non metals to lose electrons** because the amount of energy required (ionisation energy) is very large. Therefore non metals do not easily form positively charged ions. Non metals therefore **react by gaining electrons** to form negatively charged ions.

Ion formation

Halogens have seven electrons in their outermost energy level. They react by gaining one electron to attain a stable electron configuration and form negatively charged ions. During ion formation, energy is released. The energy change for this process of electron gain is called **electron affinity**.

$$F + e^- \longrightarrow F^- \text{ (Elctron affinity = –322 kJmol}^{-1}\text{)}$$

$$Cl + e^- \longrightarrow Cl^- \text{ (Electron affinity = –349 kJmol}^{-1}\text{)}$$

$$Br + e^- \longrightarrow Br^- \text{ (Electron affinity = –325 kJmol}^{-1}\text{)}$$

$$I + e^- \longrightarrow I^- \text{ (Electron affinity = –295 kJmol}^{-1}\text{)}$$

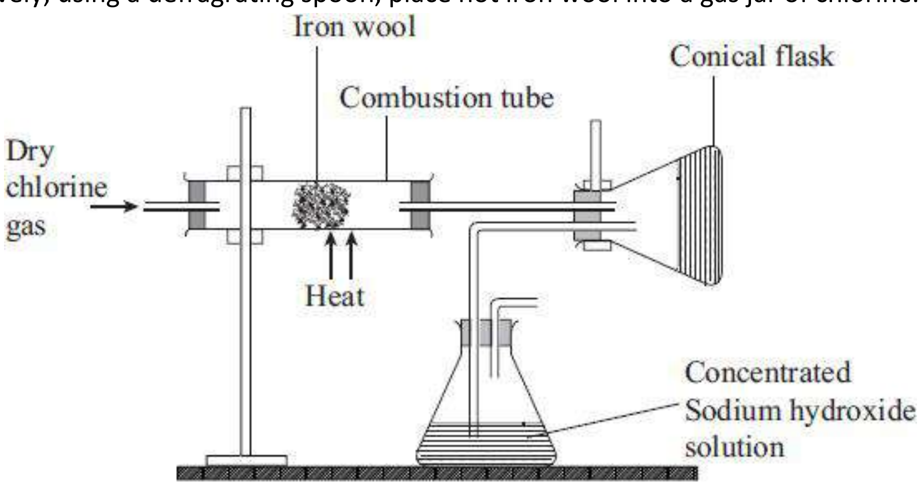
Generally, **the electron affinity decreases as the size of the atoms increases** hence **reactivity decreases down the group.**

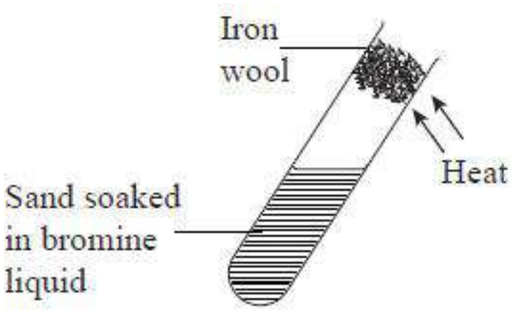
Reaction with metals

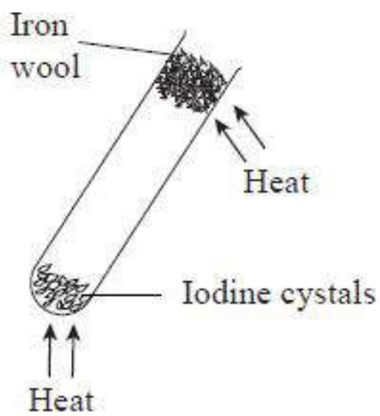
The reaction between iron and a halogen results in the formation of a salt.

The procedure below can be used to react halogens with metals. Pass a stream of dry chlorine gas over heated iron wool as shown below. Record your observations.

For bromine and iodine, heat the iron wool in a test-tube in which bromine and iodine vapour is generated and passed over the wool as shown below. The test-tube should be held with a test-tube holder. Alternatively, using a defragrating spoon, place hot iron wool into a gas jar of chlorine.







Discussion Questions

1. What is observed when the hot iron wool reacts with the halogens?

Chlorine reacts most vigorously with hot iron forming dark-brown crystals of iron (III) chloride. Hot iron glows in bromine vapour to form dark-red crystals of iron (III) bromide. Iodine vapour reacts slowly with hot iron to form grayish black crystals of iron (II) iodide. Iodine is not reactive enough to form a salt with iron.

The equations below represent the reactions which occur.

$$\text{Iron + Chlorine} \longrightarrow \text{Iron (III) chloride}$$

$$2Fe\,(s) + 3\,Cl_2(g) \longrightarrow 2FeCl_3(s)$$

$$\text{Iron + Bromine} \longrightarrow \text{Iron (III) bromide}$$

$$2Fe(s) + 3Br_2(g) \longrightarrow 2FeBr_3(s)$$

$$\text{Iron + Iodine} \longrightarrow \text{Iron (II) iodine}$$

$$Fe(s) + I_2(g) \longrightarrow FeI_2(s).$$

2. State and explain the role of the concentrated sodium hydroxide.

Concentrated sodium hydroxide is used to react with excess chlorine to avoid emitting poisonous chlorine gas into the air.

3. Write chemical equations for the reactions between chlorine, bromine and iodine with zinc.

Halogens react with heated zinc to form zinc salts.

$$\text{Zinc + Chlorine} \longrightarrow \text{Zinc chloride}$$

$$Zn(s) + Cl_2(g) \longrightarrow ZnCl_2(s)$$

$$\text{Zinc + Bromine} \longrightarrow \text{Zinc bromide}$$

$$Zn\,(s) + Br\,(g) \longrightarrow ZnBr_2(s)$$

$$\text{Zinc + Iodine} \longrightarrow \text{Zinc iodide}$$

$$Zn(s) + I_2(g) \longrightarrow ZnI_2(s)$$

Other salts formed in the same method are $MgCl_2$, $AlCl_3$ and NaCl.

4. What is the order of reactivity of halogens? Explain

The reaction between chlorine and metals is more vigorous than that of bromine. The order of reactivity of the halogens with metals **decreases down** the group.

Chlorine

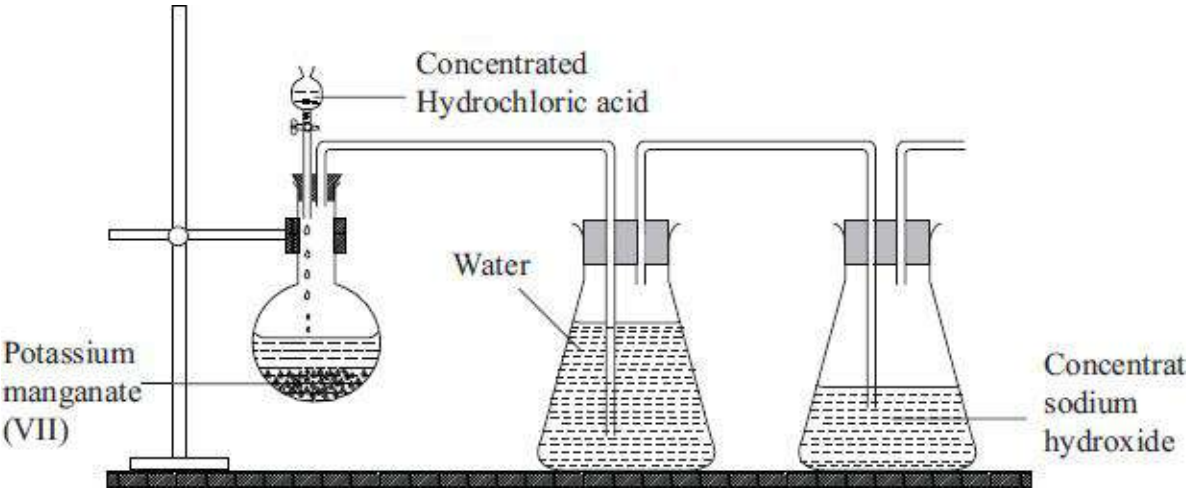
Bromine ↓ decreasing reactivity

Iodine

The ability of an atom to **gain an electron** in its outermost energy level **decreases as the size of the atoms increase,** hence the **decrease in reactivity of halogens down the group.**

Reaction with water

To study the reaction between halogens and water, bubble chlorine through water in a conical flask for a few minutes using an experimental set-up as shown below.



Observe the colour of the resulting solution. Test the solution with litmus paper.

Discussion Questions

1. What is observed when chlorine is passed through water for some time?

Chlorine dissolves in water to form chlorine water which is a mixture of hydrochloric acid and chloric (I) acid.

$$\text{Chlorine + water} \longrightarrow \text{Hydrochloric acid + Chloric (I) acid}$$

$$Cl_2(g) + H_2O(g) \longrightarrow HCl(aq) + HClO\,(aq)$$

2. What changes are observed on the litmus papers? Explain

When the chlorine water is tested with litmus paper, the blue one turns red, showing that the solution is acidic. Then the litmus papers are **bleached (decolourised) immediately.**

The bleaching action is a **property of Chloric (I) acid**. Chloric (I) acid is **unstable** and **decomposes to form hydrochloric acid** and an **atom of oxygen**. The oxygen atom then combines with the natural dye in the litmus papers to form a colourless compound.

$$\underset{(coloured)}{\text{Dye}} + HClO(aq) \longrightarrow \underset{(colourless)}{(\text{Dye} + O)} + HCl(aq)$$

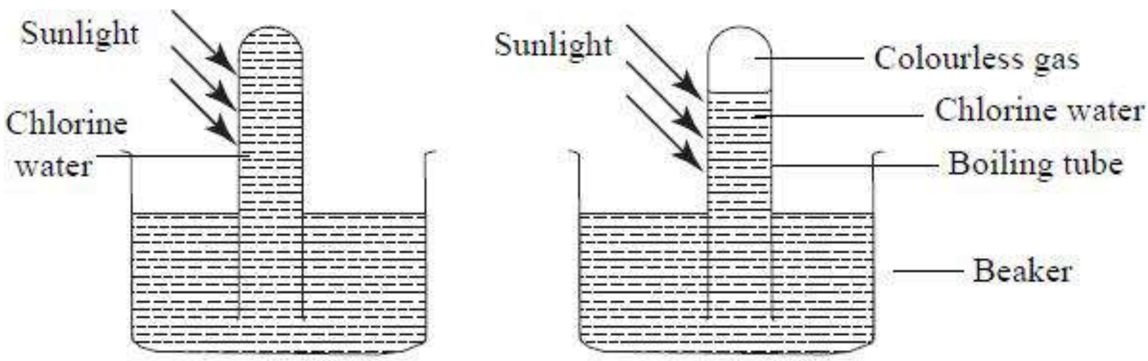
Chlorine **does not bleach dry litmus pape**r because **chloric (I) acid cannot be formed in the absence of water.** The bleaching action is only possible in the presence of water.

Chlorine water is y**ellow due to the presence of chloric (I) acid.**

In **sunlight**, the **chlorine water is decolourised** due to the decomposition of chloric (I) acid to oxygen gas and hydrochloric acid by the sunlight.

$$HClO(aq) \xrightarrow{Sunlight} 2HCl(aq) + O_2(g)$$

This reaction **does not take place in the dark**.



(a) At the beginning (b) After one day

Some uses of Halogens and their Compounds

1. Fluorine is a raw material in the preparation of a synthetic fibre known as polytetraflouroethene.
2. Some compounds of fluorine are added to water and some tooth pastes in small quantities to reduce tooth decay.
3. Fluorine is used to manufacture hydrogen fluoride used to engrave words or pictures on glass.
4. Chlorine is used to make bleaches used in paper pulp and textile industries.
5. Chlorine is added to water to kill micro-organisms in water treatment works.
6. Chlorine is used in the manufacture of a plastic known as polyvinylchloride (PVC).
7. Chlorine is used in the large-scale manufacture of hydrochloric acid.
8. Bromine is used in the manufacture of silver bromide which is used to make the light sensitive photographic paper and films.
9. A solution of iodine in alcohol (tincture of iodine) is used as a disinfectant.

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4. Noble gases

The **elements in group (VIII)** of the periodic table are called **noble gases.** Noble gases are found as free atoms in nature and form about 1% of air. They include helium, neon, argon, krypton, xenon and radon. Argon is the most abundant and forms about 0.9% of air by volume.

Noble gases were initially called **inert gases** because they were thought to be unreactive.

Task: Draw the atomic structure of the first 3 noble gases

Electron Arrangements of the first three Noble Gases

Element	Symbols	Atomic number	Electronic arrangement
Helium	He	2	2
Neon	Ne	10	2.8
Argon	Ar	18	2.8.8

Helium with only two electrons has one occupied energy level which is full. Hence it has a **duplet**. The rest have eight electrons in their outermost occupied energy level. Thus they have the **octet**.

Under normal conditions noble gases neither gain nor lose electrons. They are therefore stable and non reactive.

Physical Properties of Noble gases

Group VIII elements are colourless monoatomic gases.

Element	Symbol	Atomic number	Atomic radius (nm)	1st ionization energy ($kJmol^{-1}$)	Melting point (°C)	Boiling point (°C)
Helium	He	2	0.128	2372	–270	–269
Neon	Ne	10	0.160	2080	–249	–246
Argon	Ar	18	0.192	1520	–189	–186
Krypton	Kr	36	0.197	1350	–157	–152
Xenon	Xe	54	0.217	1170	–112	–108

Discussion Questions

1.What happens to each of the following properties down the group?

(a) Atomic radius.

Atomic radii increase down the group due to the increase in the number of energy levels. The increase in atomic radii down the group explains why the first ionization energy of the gases decreases down the group.

(b) Melting point and boiling point.

Noble gases have low melting and boiling points. This is because of the weak inter atomic forces of attraction between the atoms. However, as the atomic size increases down the group there is increase in strength of inter atomic forces of attraction between atoms. Hence the rise in melting and boiling points down the group.

(c) Ionization energies.

Helium has a duplet electron arrangement while the others have an octet in their outermost energy level. Therefore they all have a stable electron arrangement. This explains the high ionization energies for all the elements.

2.Xenon takes part in some reactions. Explain.

Xenon has a **large atomic radius.** In xenon an electron in the outermost energy level is **relatively far** from the positive nucleus. It therefore has a **low ionization energy compared to the other noble gases.** For this reason, xenon takes part in some reactions.

Uses of Noble Gases

The inert nature of the noble gases enables them to have a wide range of uses.

1. Argon is used in light bulbs to provide an inert environment to prevent oxidation.
2. Argon is used as an insulator in arch-welding.
3. Neon gas is used in street and advertisement lights.
4. Helium mixed with oxygen is used in deep sea diving and mountaineering. The mixture is also used in hospitals for patients with respiratory problems and those undergoing certain forms of surgery.
5. Helium can be used instead of hydrogen in balloons for meteorological research.
6. Helium is used in thermometers for the measurement of very low temperatures.
7. Liquid helium is used to keep certain metal alloys at temperatures low enough for them to become super conductors.

5. Properties and Trends across A period

Elements in the same period have the same number of occupied energy levels. As you move across the period, the number of electrons in the outermost energy level increases by one. While the elements in the same group exhibit similar properties, those across a period show a gradual change in properties. This can be illustrated by considering elements in Period 3 from left to right.

(a) Trends in Physical Properties of Elements in Period 3

The table below gives a summary of the physical properties of Elements in Period 3.

	Na	Mg	Al	Si	P (white)	S (monoclinic)	Cl	Ar
Physical state and appearance	Silver	Silver solid	Silver solid	Black solid	White solid	Yellow solid	Greenish yellow gas	Colour-less
Electron arrangment	2.8.1	2.8.2	2.8.3	2.8.4	2.8.5	2.8.6	2.8.7	2.8.8
Valency	1	2	3	4	3 or 5	2	1	0
Atomic radius	0.157	0.136	0.125	0.117	0.110	0.104	0.09	0.192
M.P (°C)	98	650	660	1410	44	119	–101	–189
B.P (°C)	890	1110	2470	2360	280	445	–35	–186
Structure	Giant metallic	Giant metallic	Giant metal-lic	Giant atomic	Mo-lecular	Molecular	Molecu-lar	Exists as atoms
Bond type	Metallic	Metallic	Metal-lic	Cova-lent	Cova-lent	Covalent	Covalent	Van der waals

(i) Trends in Electrical conductivity

Discussion Questions

1. List the elements in period 3 that conduct electricity.

Sodium, magnesium and aluminium are good conductors of electricity.

2.Explain the trends in the electrical conductivity of the elements in period 3.

Sodium, magnesium and aluminium **have delocalised electrons** in their structures. These delocalised electrons are responsible for the conduction of electricity. **Conductivity increases with increase in the number of delocalised electrons.** Therefore, **aluminium** with three delocalised electrons from each atom in the structure **has** the **highest electrical conductivity**.

Phosphorus, sulphur, chlorine and argon **are all made up of molecules** and therefore are **non-conductors** of electricity.

Silicon is unique among the elements because it is a **semi-conductor**. Its electrical conductivity **increases with increase in temperature.**

(ii) Other physical properties

Discussion Questions

1. **Explain the trends in atomic radii across the period**

The atomic radii of the elements **gradually decrease across the period** from left to right. This is explained by the **increase in the nuclear charge across the period** due to an increase in the number of protons. Although there is an additional number of electrons, they enter the same energy level. This means that the shielding effect remains the same as the nuclear charge increases. The forces of attraction between the nuclei of these elements and the electrons in the outermost energy levels progressively increase across the period. As a result, the electrons in the outermost energy level are pulled closer to the nucleus, thereby decreasing the size of the atoms across the period from sodium to chlorine.

2. **Explain the trends in melting points and boiling points across the period.**

Sodium, magnesium and aluminium have giant metallic structures. Therefore, they have strong metallic bonds. These bonds require a lot of energy to break hence high melting and boiling points. Silicon has giant atomic structure hence high melting and boiling points. The rest of the non-metals have molecular structures held together by weak Van der Waals forces which require little energy to break hence low melting and boiling points.

3. **Why is the melting point of aluminium higher than that of sodium?**

Aluminium contributes three electrons to the metallic lattice whereas sodium contributes only one. Also, due to the small size of the aluminium atom, the packing of the atoms is close. Therefore, the metallic bonds in aluminium are stronger than in sodium and magnesium, hence the higher melting and boiling points of aluminium.

4. **Silicon, a non-metal, has a much higher melting point than all the other elements in the period. Explain.**

Silicon has a **giant atomic structure** in which **all the atoms are held together by strong covalent bonds.** These need **a lot of heat energy to break, hence the high melting and boiling points of silicon.** In contrast, **phosphorus and chlorine are molecular.** The atoms in the molecules are held together by strong covalent bonds while **the molecules themselves are held together by van der Waals forces** which require **little energy to break**. Melting involves breaking the van der Waals forces.

5. **Why are boiling points of chlorine and argon very low?**

Chlorine and argon exist as gases at room temperature. They have low melting and boiling points due to the presence of weak van der Waals forces. Chlorine is diatomic while argon is monoatomic.

(b) Trends in Chemical Properties of the Elements in Period 3

Summary of reaction of period 3 elements with air, water and dilute acids.

Element	Na	Mg	Al	Si	P	S	Cl
Reaction with air or oxygen	Readily reacts with air. Burns brightly in oxygen to form Na_2O	Reacts slowly with air. Burns in oxygen with a bright flame to form MgO	Forms a protective layer of Al_2O_3 when it burns in oxygen	Silicon powder burns at temperatures above 950°C to form SiO_2	White phosphorus smoulders in air. Burns in air to form P_2O_3 and P_2O_5	Burns in air or oxygen to form SO_2	No reaction with air or oxygen under normal conditions
Reaction with water	Vigorous reaction evolving H_2 and NaOH	Slow reaction with cold water but reacts rapidly with steam.	No reaction with water	No reaction	No reaction	No reaction	Dissolves in water to form chlorine water
Reaction with dilute acids	Violent reaction giving H and a salt.	Rapid evolution of H and a salt formed.	Reacts slowly to form H_2	No reaction	No reaction	No reaction	No reaction

Reaction with oxygen

Discussion Questions

1. **Describe how elements in Period 3 react with oxygen. Explain the nature of the resulting solutions when the products are dissolved in water.**

- Sodium **reacts vigorously with oxygen** to form a white solid, sodium oxide.

  $$\text{Sodium} + \text{Oxygen} \longrightarrow \text{Sodium oxide}$$
  $$4Na(s) + O_2\,(g) \longrightarrow 2Na_2O\,(s)$$

  The sodium oxide produced in the reaction readily dissolves in water to form an alkaline solution

  $$\text{Sodium oxide} + \text{Water} \longrightarrow \text{Sodium hydroxide}$$
  $$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

- Magnesium **burns with a bright white light** to give a white solid, magnesium oxide.

  $$\text{Magnesium} + \text{Oxygen} \longrightarrow \text{Magnesium oxide}$$
  $$2Mg\,(s) + O_2(g) \longrightarrow 2MgO(s)$$

  The magnesium oxide produced in the reaction is slightly soluble in water. The solution formed is alkaline.

  $$\text{Magnesium oxide} + \text{water} \longrightarrow \text{Magnesium hydroxide}$$
  $$MgO(s) + H_2O(l) \longrightarrow Mg(OH)_2(aq)$$

- Aluminium foil is usually coated with **a layer of aluminium oxide, Al_2O_3.** This **prevents the reaction with oxygen.** When polished, **aluminum reacts slowly** with oxygen to form the white solid, aluminium oxide.

  $$\text{Aluminium} + \text{Oxygen} \longrightarrow \text{Aluminium oxide}$$
  $$4Al(s) + 3O_2(g) \longrightarrow 2Al_2\,O_3(s)$$

  The aluminium oxide is insoluble in water.

- Silicon powder **can only burn in oxygen at high temperatures (about 450°C)** to form solid silicon (IV) oxide.

$$\text{Silicon + Oxygen} \longrightarrow \text{Silicon (IV) oxide}$$

$$Si(s) + O_2(g) \longrightarrow SiO_2(s)$$

Silicon (IV) oxide is insoluble in water.

- Phosphorus **readily burns in oxygen with a bright orange flame** to form a white solid, phosphorus (V) oxide.

$$\text{Phosphorus + Oxygen} \longrightarrow \text{Phosphorus (V) oxide}$$

$$P_4(s) + 5O_2(g) \longrightarrow 2P_2O_5(s)$$

Phosphorus (V) oxide readily dissolves in water to form an acidic solution.

$$\text{Phosphorus (V) oxide + Water} \longrightarrow \text{Phosphorus (V) acid}$$

$$P_2O_5(s) + 3H_2O(l) \longrightarrow 2H_3PO_4(aq)$$

- Sulphur **burns in oxygen with a blue fame** to form a gas sulphur (IV) oxide.

$$\text{Sulphur + Oxygen} \longrightarrow \text{Sulphur (IV) oxide}$$

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

The sulphur (IV) oxide gas readily dissolves in water to give an acidic solution of sulphuric (IV) acid, H_2SO_3, which is easily oxidized to sulphuric (VI) acid H_2SO_4.

$$\text{Sulphur (IV) oxide + Water} \longrightarrow \text{Sulphuric (IV) acid}$$

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

$$\text{Sulphuric (IV) acid + Oxygen} \longrightarrow \text{Sulphuric(VI) acid}$$

$$2H_2SO_3(aq) + O_2(g) \longrightarrow 2H_2SO_4(aq)$$

- Chlorine burns in oxygen under certain conditions to form **acidic oxides** while argon is not reactive.

2. State the trend in how elements in period 3 react with oxygen

The following trends in the elements of period 3 can be identified.

- **All** the elements across period 3**,** with the **exception of argon**, **burn** in oxygen to form oxides.
- The **reactivity of the metals with oxygen decreases from left to right** across the period. Sodium is the most reactive of the three metals in the period and aluminium the least. The order of reactivity with oxygen is therefore Na > Mg > Al. This is **because of the increase in nuclear charge from sodium to aluminium, which makes it easier to remove an electron from a sodium atom than from an aluminium atom.**
- **Metallic elements** burn in oxygen to form **basic oxides. Soluble** metallic oxides dissolve in water to form **alkaline solutions.**
- The **reactivity of the non-metallic elements with oxygen increases from left to right across the third period.** This is **because the ease of gaining electrons increases from left to right.** Phosphorus is the least reactive and chlorine the most reactive. The order of reactivity with oxygen, starting with the most reactive is therefore: Cl > S > P. **Non-metals react by gaining electrons.**
- The **non-metallic elements** burn in oxygen to form **acidic oxides** which dissolve in water to form **acidic solutions**.

Reaction with Water

Discussion Questions

1. Describe how each of the elements in period 3 react with water.

- Sodium **reacts violently with cold water** to form sodium hydroxide and hydrogen gas.

Sodium + Water ⟶ Sodium hydroxide + Hydrogen

$$2Na(s) + H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

- Magnesium **reacts very slowly with cold water** to form magnesium hydroxide and hydrogen gas.

Magnesium + water ⟶ Magnesium hydroxide + Hydrogen

$$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

- Aluminium **does not normally react with cold water or steam** due to the presence of a **coating of aluminium oxide, which prevents any reaction.** However, **at temperatures above 700 °C steam can react with aluminium.** Due to its apparent inability to react with water, aluminium was prefered for making cooking vessels such as ***pans*** and ***sufurias*.**

2. **Explain why the reactivity varies across the period.**

- In general, the reactivity of the metals with water **decreases** from sodium to aluminium, Na > Mg > Al. sodium is **more reactive** becaus**e it loses its valence electron more readily** than magnesium and aluminium.
- Non-metals **do not displace hydrogen from water**. Therefore, silicon, phosphorus and sulphur **do not react with either cold water or steam.**
- Chlorine is an exception since **it dissolves in water** to form **chlorine water**, which is a **mixture of hydrochloric acid and chloric (I) acid.**

Chlorine + water ⟶ Hydrochloric acid + Hypochlorous acid

$$Cl_2(g) + H_2O(l) \longrightarrow HCl(aq) + HOCl(aq)$$

Reaction with Acids

Discussion Questions

Explain the trend in reactivity of the period 3 elements with dilute acids.

- Magnesium reacts with both dilute hydrochloric acid and dilute sulphuric (VI) acid to form a salt and hydrogen gas.

Magnesium + hydrochloric acid ⟶ Magnesium chloride + Hydrogen gas

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

Magnesium + Sulphuric (VI) acid ⟶ Magnesium sulphate + Hydrogen gas.

$$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$$

- Aluminium does not readily react with dilute acid due to the **presence of an aluminium oxide coating.** However, on removing the oxide coating, the metal reacts.

Aluminium + Hydrochloric acid ⟶ Aluminium chloride + Hydrogen gas

$$2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2(g)$$

$$2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO)_4(aq) + 3H_2(g).$$

- The reaction between sodium and acids is explosive and should NEVER be tried. The order of reactivity with acid is Na > Mg > Al.
- Silicon, phosphorus, sulphur and chlorine **do not react with dilute acids.**

1. 2006 Q19 P1

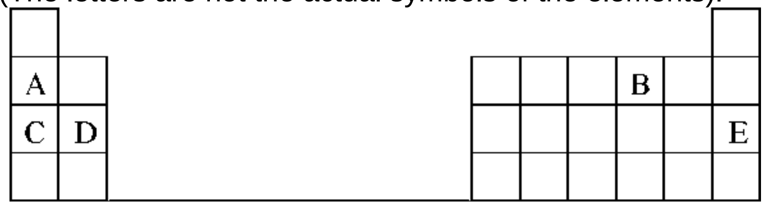
(a) Starting from solid magnesium oxide, describe how a solid sample of magnesium hydroxide can be prepared. (2 marks)

(b) Give one use of magnesium hydroxide. (1 mark)

2. 2006 Q 3a,b P2

(a) Distinguish between isotopes and allotropes.

(b) The chart below is part of the periodic table. Study it and answer the questions that follow. (The letters are not the actual symbols of the elements).



(i) Select the element in period three which has the shortest atomic radius. Give a reason for you answer. (2 marks)

(ii) Element F has the electronic structure, 2.8.18.4. On the chart above; indicate the position of element F. (1 mark)

(iii) State one use of the elements of which E is a member. (1 mark)

3. 2007 Q3a P1

Both chlorine and iodine are halogens. What are halogens? (1 mark)

4. 2007 Q 8 P1

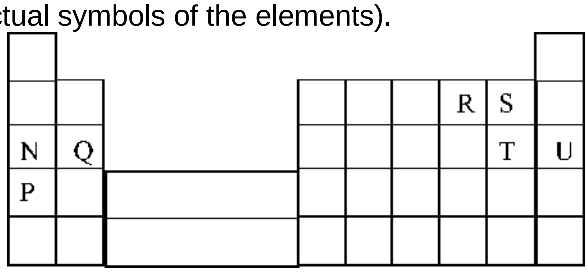
Explain why there is general increase in the first ionization energies of the elements in period 3 of the periodic table from left to right.
(2 marks)

5. 2007 Q 25b

Give a reason why helium is increasingly being preferred to hydrogen in weather balloons. (1 mark)

6. 2008 Q 18 P1

The grid below is part of the periodic table. Use it to answer the questions that follow, (the letters are not the actual symbols of the elements).



(a) Indicate on the grid the position of an element represented by letter V whose atomic number is 14. (1 mark)

(b) Select a letter which represents a monoatomic gas. (1 mark)

7. 2009 Q 1 P1

The ionization energies for three elements A, B and C are shown in the table below.

Element	**A**	**B**	**C**
Ionisation energy (kJ /mole)	519	418	494

(a) What is meant by ionization energy? (1 mark)

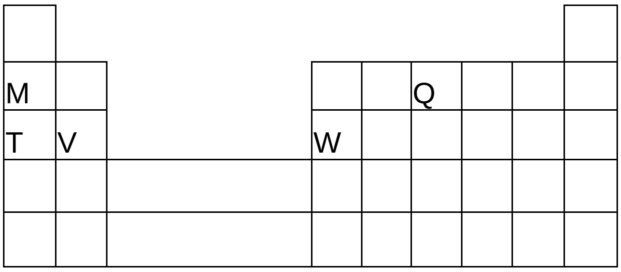
(b) Which element is the strongest reducing agent? Give a reason (2 marks)

8. 2009 Q 21 P1,2011 Q 5 P1

Give the name of the product formed when magnesium reacts with phosphorus. (1 mark)

9. 2011 Q5 P1

The diagram below represents part of the periodic table. Use it to answer the questions that follow.



(a) Write the electron arrangement for the stable ion formed by W (1 mark)
(b) Write an equation for the reaction between V and Q (1 mark)
(c) How do the ionization energies of the elements M and T compare? (1 mark)

10. 2011 Q 22 P1

The table below gives some properties of three elements in group (VII) of the periodic table. Study it and answer the questions that follow.

Element	**Atomic No.**	**Melting Point (°C)**	**Boiling Point (°C)**
Chlorine	17	-101	-34.7
Bromine	35	-7	58.8
Iodine	53	114	184

(a) Which element is in liquid form at room temperature? Give a reason. (1 mark)
(b) Explain why the boiling point of iodine is much higher than that of chlorine. (2 marks)

11. 2011 Q 31 P1

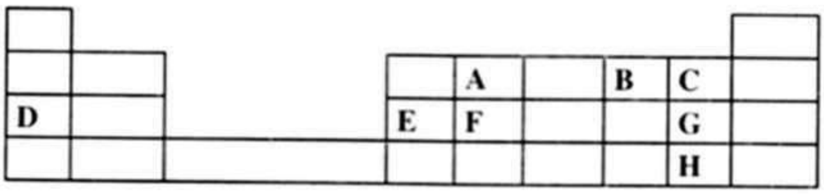
What name is given to elements which appear in group (II) of the periodic table? (1 mark)

12. 2012 Q14 P1

Distinguish between ionisation energy and electron affinity of an element. (2 marks)

13. 2012 Q2 P2

The grid below is part of the periodic table. Use it to answer the questions that follow. (the letters are not the actual symbols of the elements).



(a) Which is the most reactive non-metallic element shown in the table? Explain. (2 marks)

(b) (i) Write the formula of the compound formed when element A reacts with element B (1 mark)
(ii) Name the bond type in the compound formed in b (i) above (1 mark)

(c) (i) What is the name given to the group of elements where, C, G and H belong?
(ii) Write an equation for the reaction that occurs when C in gaseous form is passed through a solution containing ions of element H (1 mark)

(d) The melting points of elements F and G are 14100C and -101 respectively. In terms of structure and bonding, explain why there is a large difference in the melting points of F and G.

(2 marks)

(e) D forms two oxides. Write the formula of each of the two oxides. (1 mark)

(f) J is an element that belongs to the 3rd period of the periodic table and a member of the alkaline earth elements. Show the position of j in the grid (1 mark)

14. 2012 Q4 P1

The table below shows properties of some elements A, B, C and D which belong to the same period of the periodic table. The letters are not the actual symbols of the elements.

Element	**A**	**B**	**C**	**D**
MP (°C)	1410	98	-101	660
Atomic radii (nm)	0.117	0.186	0.099	0.143
Electrical conductivity	Poor	Good	Non-conductor	Good

(a) Arrange the elements in the order they would appear in the period. Give a reason (2 marks)

(b) Select the metallic element which is the better conductor of electricity. Give a Reason. (1 mark)

15. 2012 Q27 P1

(a) The electronic arrangement of the ion of element Q is 2.8.8. If the Formula of the ion is Q^{3-}, state the group and period to which Q belongs,

Group: (½ mark)

Period: (½mark)

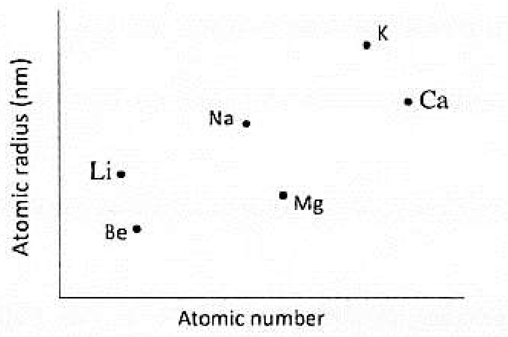
(b) Helium, neon and argon belong to group 8 of the periodic table. Give:

(i) The general name of these elements; (1 mark)

(ii) One use of these elements (1 mark)

16. 2013 Q20 P1

The plots below were obtained when the atomic radii of some elements in groups I and II were plotted against atomic numbers.



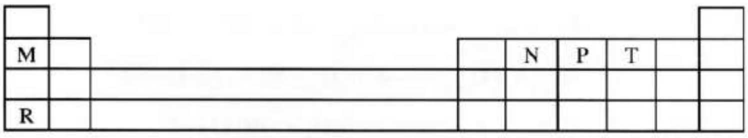
Explain:

(a) The trend shown by Li, Na and K. (1 mark)

(b) Why the atomic radii of elements Be, Mg, and Ca are lower than those of Li, Na and K. (2 marks)

17. 2013 Q1 P2

(a) The grid below represents part of the periodic table. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.



(i) Select a letter which represents an element that losses electrons most readily. Give a reason for your answer. (2 marks)

(ii) Explain why the atomic radius of P is found to be smaller than that of N. (2 marks)

(iii) Element M reacts with water at room temperature to produce 0.2 dm^3 of gas. Determine the mass of M which was reacted with water. (molar gas volume at r.t.p is 24 dm^3, relative atomic mass of M = 7). (3 marks)

(b) Use the information in the table below to answer the questions that follow. (The letters are not the actual symbols of the elements).

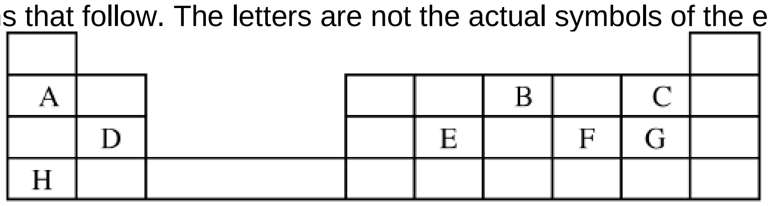
Element	State of oxide at room temperature	Type of oxide	Bonding in oxide
U	Solid	Acidic	Covalent
W	Solid	Basic	Ionic
X	Liquid	Neutral	Covalent
Y	Gas	Neutral	Covalent

Identify a letter which represents an element in the table that could be calcium, carbon or sulphur. Give a reason in each case.

(i) Calcium_____________ (½ mark)
Reason_____________ (½ mark)
(ii) Carbon _____________ (½ mark)
Reason______________ (½ mark)
(iii) Sulphur _____________ (½ mark)
Reason______________ (½ mark)

18. 2014 Q2 P1

(a) The grid below represents part of the periodic table. Study it and answer the questions that follow. The letters are not the actual symbols of the elements



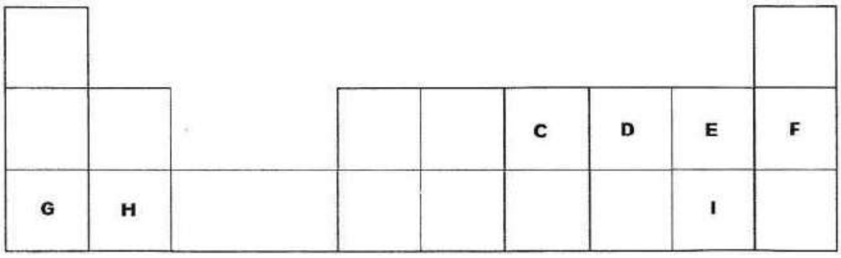
(i) Select the most reactive metal. Explain (2 marks)
(ii) Select an element that can form an ion with a charge of 3^- (1 mark)
(iii) Select an alkaline earth metal (1 mark)
(iv) Which group 1 element has the highest first ionization energy? Explain (2 marks)
(v) Element A combines with chlorine to form a chloride of A. State the most likely pH value of a solution of a chloride of A. Explain (2 marks)

(b) (i) Explain why molten calcium chloride and magnesium chloride conduct electricity while carbon tetrachloride and silicon tetrachloride do not. (2 marks)

(ii)Under the same conditions, gaseous neon was found to diffuse faster than gaseous fluorine. Explain this observation. (F=19.0; Ne=20.0)
(2 marks)

19. 2015 Q23 P1

The table below is part of the periodic table. The letters are not the actual symbols of the elements. Study it and answer the questions that follows.

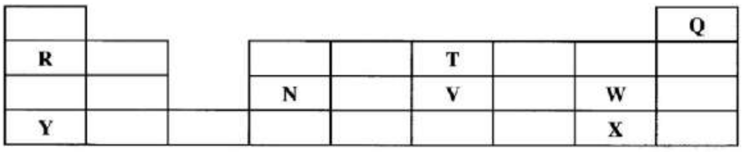


(a) Select an element which is stored in paraffin in the laboratory. (1 mark)

(b) How do the ionic radii of E and I compare? Explain (2 marks)

20. 2015 Q3c-d P2

(a) The diagram below shows part of a periodic table. The letters do not represent the actual symbols of elements. Use the diagram to answer the questions that follow.



(i) Explain why the oxidizing power of W is more than that of X. (1 mark)
(ii) How do the melting points of R and T compare? Explain. (2 marks)

(iii) Sketch an element that could be used
II. In weather balloons (1 mark)
III. For making a cooking pot (1 mark)

(b) (i) Classify the substances water, iodine, diamond and candle wax into elements and compounds (2 marks)

Elements	Compounds

(i) Give one use of diamond. (1 mark)

21. 2015 Q20 P1

A crystal of iodine, heated gently in a test tube gave off a purple vapour.

(a) Write the formula of the substance responsible for the purple vapour. (1 mark)

(b) What type of bond is broken when the iodine crystal is heated gently? (1 mark)

(c) State one use of iodine. (1 mark)

22. 2016 Q1 P2
Use the information in the table below to answer the questions that follow. The letters do not represent the actual symbols of the elements.

Element	Atomic Number	Melting Point (°C)
R	11	97.8
S	12	650.0
T	15	44.0
U	17	-102
V	18	-189
w	19	64.0

(a) Give reasons why the melting point of:
(i) S is higher than that of R; (1 mark)
(ii) V is lower than that of U. (2 marks)

(b) How does the reactivity of W with Chlorine compare with that of R with chlorine? Explain.
(2 marks)
(c) Write an equation for the reaction between T and excess oxygen.
(1 mark)
(d) When 1.15g of R were reacted with water, 600cm3 of gas was produced. Determine the relative atomic mass of R. (Molar gas volume = 24 000 cm^3). (3 marks)

(e) Give one use of element V. (1 mark)

23. 2017 Q1 P1
Table 1 shows the atomic numbers and the first ionisation energies of three elements. The letters are not actual symbols of the elements. Use it to answer the questions that follow.

Table 1

Element	Atomic number	First ionisation energy kJmol⁻
A	3	519
B	11	494
C	19	418

(a) Explain the trend in first ionisation energy from A to C. (2 mark)
(b) Write the electronic configuration for the ion of C. (1 mark)

24. 2018 P1 Q5.
Describe an experiment to show that group one elements react with cold water to form alkaline solutions. (3 marks)

25. 2018 Q2 P2

Figure 2 is a section of the periodic table. Study it and answer the questions that follow. The letters do not represent the actual symbols of elements

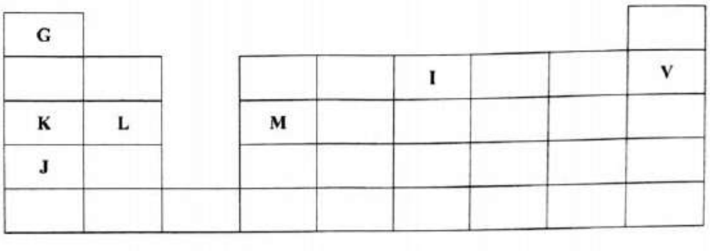


Figure 2

(a) (i) Select elements which belong to the same chemical family. (1 mark)
(ii) Write the formulae of ions for elements in the same period. (1 mark)

(b) The first ionisation energies of two elements K and M at random are 577 kJ/mol and 494 kJ/mol. (1 mark)
(i) Write equations for the 1st ionisation energies for elements K and M and indicate their energies. (1 mark)
(ii) Explain the answer in b(i). (1 mark)
(iii) Write the formula of the compound formed when **L** and **I** react. (1 mark)
(iv) Give one use of element **V.** (1 mark)

(c) (i) State another group that **G** can be placed in **Figure 1**. Explain. (2 marks)
(ii) How do the reactivity of elements **J** and **K** compare? Explain. (2 marks)

(d) (i) Elements **L** and **M** form chlorides. Complete the following table by writing the formulae of each chloride and state the nature of the solutions. (2 marks)

Element	**Formula of chloride**	**Nature of chloride solution**
L		
M		

(ii) The chloride of element M vaporises easily while its oxide has a high melting point. Explain. (2 marks)

26. 2019 P1 Q21

Study the information in Table 3 and use it to answer the questions that follow.

Elements	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
Atomic Numbers	11	12	13	14	15	16	17
Atomic radii(nm)	0.157	0.136	0.125	0.117	0.110	0.104	0.099

(a) Explain the trend in atomic radii from sodium to chlorine. (1 mark)

(b) Explain how the chloride of aluminium differs from those of other metals in the period. (2 marks)