CHEMICAL FAMILIES (PATTERNS AND PROPERTIES)

The number of valence electrons and the number of occupied energy levels in an atom of an element determine the position of an element in the periodic table.i.e The number of occupied energy levels determine the Period and the valence electrons determine the Group.
Elements in the same group have similar physical and chemical properties. The trends in physical and chemical properties of elements in the same group vary down the group. Elements in the same group thus constitute a chemical family.

(a) Group I elements: Alkali metals

Group I elements are called **Alkali metals** except Hydrogen which is a non metal. The alkali metals include:

Element	Symbol	Atomic number	Electron structure	Oxidation state	Valency
Lithium	Li	3	2:**1**	Li^+	1
Sodium	Na	11	2:8:**1**	Na^+	1
Potassium	K	19	2:8:8:**1**	K^+	1
Rubidium	Rb	37	2:8:18:8:**1**	Rb^+	1
Caesium	Cs	55	2:8:18:18:8:**1**	Cs^+	1
Francium	Fr	87	2:8:18:32:18:8:**1**	Fr^+	1

All alkali metals atom has one electron in the outer energy level. They therefore are **monovalent**. They donate /lose the outer electron to have oxidation state M^+
The number of energy levels increases down the group from Lithium to Francium.
The more the number of energy levels the bigger/larger the atomic size. e.g.
The atomic size of Potassium is bigger/larger than that of sodium because Potassium has more/4 energy levels than sodium (3 energy levels).

Atomic and ionic radius
The distance between the centre of the nucleus of an **atom** and the outermost energy level occupied by electron/s is called **atomic radius.** Atomic radius is measured in **nanometers**(n).The higher /bigger the atomic radius the bigger /larger the atomic size.

The distance between the centre of the nucleus of an **ion** and the outermost energy

level occupied by electron/s is called **ionic radius.** Ionic radius is also measured in **nanometers**(n).The higher /bigger the ionic radius the bigger /larger the size of the ion.
Atomic radius and ionic radius depend on the number of energy levels occupied by electrons. The more the number of energy levels the bigger/larger the atomic /ionic radius. e.g.
The atomic radius of Francium is bigger/larger than that of sodium because Francium has more/7 energy levels than sodium (3 energy levels).
Atomic radius and ionic radius of alkali metals increase down the group as the number of energy levels increases.

The atomic radius of alkali metals is bigger than the ionic radius. This is because alkali metals react by losing/donating the outer electron and hence lose the outer energy level.

Table showing the atomic and ionic radius of some 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

The atomic radius of sodium is 0.157nM .The ionic radius of Na^+ is 0.095nM. This is because sodium reacts by donating/losing the outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the atomic radius.

Electropositivity
The ease of donating/losing electrons is called electropositivity. All alkali metals are electropositive. Electropositivity increase as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The outer electrons experience less nuclear attraction and can be lost/ donated easily/with ease. Francium is the most electropositive element in the periodic table because it has the highest/biggest atomic radius.

Ionization energy
The minimum amount of energy required to remove an electron from an atom of element in its gaseous state is called **1^{st} ionization energy**. The SI unit of ionization energy is **kilojoules per mole/kJmole^{-1} .**Ionization energy depend on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer electrons/energy level and thus the lower the ionization energy. For alkali metals the 1^{st} ionization energy decrease down the group as the atomic radius increase

and the effective nuclear attraction on outer energy level electrons decrease.
e.g. The 1st ionization energy of sodium is 496 kJmole^{-1} while that of potassium is 419 kJmole^{-1} .This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from sodium to Potassium. It requires therefore less energy to donate/lose outer electrons in Potassium than in sodium.

Physical properties

Soft/Easy to cut: Alkali metals are soft and easy to cut with a knife. The softness and ease of cutting increase down the group from Lithium to Francium. This is because an increase in atomic radius, decreases the strength of metallic bond and the packing of the metallic structure

Appearance: Alkali metals have a shiny grey metallic luster when freshly cut. The surface rapidly/quickly tarnishes on exposure to air. This is because the metal surface rapidly/quickly reacts with elements of air/oxygen.

Melting and boiling points: Alkali metals have a relatively low melting/boiling point than common metals like Iron. This is because alkali metals use only one delocalized electron to form a weak metallic bond/structure.

Electrical/thermal conductivity: Alkali metals are good thermal and electrical conductors. Metals conduct using the outer mobile delocalized electrons. The delocalized electrons move randomly within the metallic structure.

Summary of some physical properties of the 1st three alkali metals

Alkali metal	Appearance	Ease of cutting	Melting point (^oC)	Boiling point (^oC)	Conductivity	1st ionization energy
Lithium	Silvery white	Not easy	180	1330	Good	520
Sodium	Shiny grey	Easy	98	890	Good	496
Potassium	Shiny grey	Very easy	64	774	Good	419

Chemical properties

(i)Reaction with air/oxygen

On exposure to air, alkali metals reacts with the elements in the air.
Example
On exposure to air, Sodium first reacts with Oxygen to form sodium oxide.

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

The sodium oxide formed further reacts with water/moisture in the air to form sodium hydroxide solution.

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

Sodium hydroxide solution reacts with carbon(IV)oxide in the air to form sodium carbonate.

$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(g) + H_2O(l)$

(ii)Burning in air/oxygen

Lithium burns in air with a **crimson**/deep red flame to form Lithium oxide

$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$

Sodium burns in <u>air</u> with a **yellow** flame to form sodium oxide

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

Sodium burns in <u>oxygen</u> with a **yellow** flame to form sodium peroxide

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

Potassium burns in <u>air</u> with a **lilac/purple** flame to form potassium oxide

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

(iii) Reaction with water:

<u>Experiment</u>

Measure 500 cm3 of water into a beaker.
Put three drops of phenolphthalein indicator.
Put about 0.5g of Lithium metal into the beaker.
Determine the pH of final product
Repeat the experiment using about 0.1 g of Sodium and Potassium.
Caution: Keep a distance

<u>Observations</u>

Alkali metal	Observations	Comparative speed/rate of the reaction
Lithium	-Metal floats in water -rapid effervescence/fizzing/bubbling -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Moderately vigorous
Sodium	-Metal floats in water -very rapid effervescence /fizzing /bubbling -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound)	Very vigorous

	-resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	
Potassium	-Metal floats in water -explosive effervescence /fizzing /bubbling -colourless gas produced (that extinguishes burning splint with explosion /"pop" sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Explosive/burst into flames

Explanation

Alkali metals are less dense than water. They therefore float in water.They react with water to form a strongly alkaline solution of their hydroxides and producing hydrogen gas. The rate of this reaction increase down the group. i.e. Potassium is more reactive than sodium .Sodium is more reactive than Lithium.

The reactivity increases as electropositivity increases of the alkali increases. This is because as the atomic radius increases , the ease of donating/losing outer electron increase during chemical reactions.

Chemical equations

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

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

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

$2Rb(s) + 2H_2O(l) \rightarrow 2RbOH(aq) + H_2(g)$

$2Cs(s) + 2H_2O(l) \rightarrow 2CsOH(aq) + H_2(g)$

$2Fr(s) + 2H_2O(l) \rightarrow 2FrOH(aq) + H_2(g)$

Reactivity increase down the group

(iv) Reaction with chlorine:

Experiment

Cut about 0.5g of sodium into a deflagrating spoon with a lid cover. Introduce it on a Bunsen flame until it catches fire. Quickly and carefully lower it into a gas jar containing dry chlorine to cover the gas jar.

Repeat with about 0.5g of Lithium.

Caution: This experiment should be done in fume chamber because chlorine is poisonous /toxic.

Observation

Sodium metal continues to burn with a yellow flame forming white solid/fumes.

Lithium metal continues to burn with a crimson flame forming white solid / fumes. Alkali metal react with chlorine gas to form the corresponding metal chlorides. The reactivity increase as electropositivity increase down the group from Lithium to Francium.The ease of donating/losing the outer electrons increase as the atomic radius increase and the outer electron is less attracted to the nucleus.

Chemical equations

$2Li(s) + Cl_2(g) \rightarrow 2LiCl(s)$
$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$
$2Rb(s) + Cl_2(g) \rightarrow 2RbCl(s)$
$2Cs(s) + Cl_2(g) \rightarrow 2CsCl(s)$
$2Fr(s) + Cl_2(g) \rightarrow 2FrCl(s)$

Reactivity increase down the group

The table below shows some compounds of the 1st three alkali metals

	Lithium	sodium	Potassium
Hydroxide	LiOH	NaOH	KOH
Oxide	Li_2O	Na_2O	K_2O
Sulphide	Li_2S	Na_2S	K_2S
Chloride	LiCl	NaCl	KCl
Carbonate	Li_2CO_3	Na_2CO_3	K_2CO_3
Nitrate(V)	$LiNO_3$	$NaNO_3$	KNO_3
Nitrate(III)	-	$NaNO_2$	KNO_2
Sulphate(VI)	Li_2SO_4	Na_2SO_4	K_2SO_4
Sulphate(IV)	-	Na_2SO_3	K_2SO_3
Hydrogen carbonate	-	$NaHCO_3$	$KHCO_3$
Hydrogen sulphate(VI)	-	$NaHSO_4$	$KHSO_4$
Hydrogen sulphate(IV)	-	$NaHSO_3$	$KHSO_3$
Phosphate	-	Na_3PO_4	K_3PO_4
Manganate(VI)	-	$NaMnO_4$	$KMnO_4$
Dichromate(VI)	-	$Na_2Cr_2O_7$	$K_2Cr_2O_7$
Chromate(VI)	-	Na_2CrO_4	K_2CrO_4

Some **uses** of alkali metals include:

(i)Sodium is used in making sodium cyanide for extracting gold from gold ore.
(ii)Sodium chloride is used in seasoning food.
(iii)Molten mixture of sodium and potassium is used as coolant in nuclear reactors.
(iv)Sodium is used in making sodium hydroxide used in making soapy and soapless detergents.
(v)Sodium is used as a reducing agent for the extraction of titanium from Titanium(IV)chloride.

(vi)Lithium is used in making special high strength glasses
(vii)Lithium compounds are used to make dry cells in mobile phones and computer laptops.

Group II elements: Alkaline earth metals

Group II elements are called **Alkaline earth metals** . The alkaline earth metals include:

Element	Symbol	Atomic number	Electron structure	Oxidation state	Valency
Beryllium	Be	4	2:**2**	Be^{2+}	2
Magnesium	Mg	12	2:8:**2**	Mg^{2+}	2
Calcium	Ca	20	2:8:8:**2**	Ca^{2+}	2
Strontium	Sr	38	2:8:18:8:**2**	Sr^{2+}	2
Barium	Ba	56	2:8:18:18:8:**2**	Ba^{2+}	2
Radium	Ra	88	2:8:18:32:18:8:**2**	Ra^{2+}	2

All alkaline earth metal atoms have two electrons in the outer energy level. They therefore are **divalent**. They donate /lose the two outer electrons to have oxidation state M^{2+}
The number of energy levels increases down the group from Beryllium to Radium. The more the number of energy levels the bigger/larger the atomic size. e.g.
The atomic size/radius of Calcium is bigger/larger than that of Magnesium because Calcium has more/4 energy levels than Magnesium (3 energy levels).

Atomic radius and ionic radius of alkaline earth metals increase down the group as the number of energy levels increases.

The atomic radius of alkaline earth metals is bigger than the ionic radius. This is because they react by losing/donating the two outer electrons and hence lose the outer energy level.

Table showing the atomic and ionic radius of the 1st three alkaline earth metals

Element	Symbol	Atomic number	Atomic radius(nM)	Ionic radius(nM)
Beryllium	Be	4	0.089	0.031
Magnesium	Mg	12	0.136	0.065
Calcium	Ca	20	0.174	0.099

The atomic radius of Magnesium is 0.136nM .The ionic radius of Mg^{2+} is 0.065nM. This is because Magnesium reacts by donating/losing the two outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the

atomic radius.

Electropositivity
All alkaline earth metals are also electropositive like alkali metals. The electropositivity increase with increase in atomic radius/size. Calcium is more electropositive than Magnesium. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The two outer electrons in calcium experience less nuclear attraction and can be lost/ donated easily/with ease because of the higher/bigger atomic radius.

Ionization energy
For alkaline earth metals the 1^{st} ionization energy decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease.
e.g. The 1^{st} ionization energy of Magnesium is 900 $kJmole^{-1}$ while that of Calcium is 590 $kJmole^{-1}$ .This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from magnesium to calcium.
It requires therefore less energy to donate/lose outer electron in calcium than in magnesium.

The minimum amount of energy required to remove a second electron from an ion of an element in its gaseous state is called the **2^{nd} ionization energy**.
The 2^{nd} ionization energy is always higher /bigger than the 1^{st} ionization energy. This because once an electron is donated /lost form an atom, the overall effective nuclear attraction on the remaining electrons/energy level increase. Removing a second electron from the ion require therefore more energy than the first electron.
The atomic radius of alkali metals is higher/bigger than that of alkaline earth metals.This is because across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level. Increase in nuclear charge increases the effective nuclear attraction on the outer energy level which pulls it closer to the nucleus. e.g.
Atomic radius of Sodium (0.157nM) is higher than that of Magnesium (0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus.

Physical properties

Soft/Easy to cut: Alkaline earth metals are **not** soft and easy to cut with a knife like alkali metals. This is because of the decrease in atomic radius of corresponding alkaline earth metal, increases the strength of metallic bond and the packing of the metallic structure. Alkaline earth metals are

(i)ductile(able to form wire/thin long rods)
(ii)malleable(able to be hammered into sheet/long thin plates)
(iii)have high tensile strength(able to be coiled without breaking/ not brittle/withstand stress)

Appearance: Alkali earth metals have a shiny grey metallic luster when their surface is freshly polished /scrubbed. The surface slowly tarnishes on exposure to air. This is because the metal surface slowly undergoes oxidation to form an oxide. This oxide layer should be removed before using the alkaline earth metals.

Melting and boiling points: Alkaline earth metals have a relatively high melting/ boiling point than alkali metals. This is because alkali metals use only one delocalized electron to form a weaker metallic bond/structure. Alkaline earth metals use two delocalized electrons to form a stronger metallic bond /structure. The melting and boiling points decrease down the group as the atomic radius/size increase reducing the strength of metallic bond and packing of the metallic structure. e.g.

Beryllium has a melting point of 1280^{o}C. Magnesium has a melting point of 650^{o}C.Beryllium has a smaller atomic radius/size than magnesium .The strength of metallic bond and packing of the metallic structure is thus stronger in beryllium.

Electrical/thermal conductivity: Alkaline earth metals are good thermal and electrical conductors. The two delocalized valence electrons move randomly within the metallic structure.
Electrical conductivity increase down the group as the atomic radius/size increase making the delocalized outer electrons less attracted to nucleus. Alkaline earth metals are better thermal and electrical conductors than alkali metals because they have more/two outer delocalized electrons.e.g.

Magnesium is a better conductor than sodium because it has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

Calcium is a better conductor than magnesium.

Calcium has bigger/larger atomic radius than magnesium because the delocalized electrons are less attracted to the nucleus of calcium and thus more free /mobile and thus better the electrical conductor

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Summary of some physical properties of the 1st three alkaline earth metals

Alkaline earth metal	Appearance	Ease of cutting	Melting point (^{o}C)	Boiling point (^{o}C)	Conduct- ivity	1st ionization energy	2nd ionization energy
Beryllium	Shiny grey	Not	1280	3450	Good	900	1800

		easy					
Magnesium	Shiny grey	Not Easy	650	1110	Good	736	1450
calcium	Shiny grey	Not easy	850	1140	Good	590	970

Chemical properties

(i)Reaction with air/oxygen

On exposure to air, the surface of alkaline earth metals is slowly oxidized to its oxide on prolonged exposure to air.

Example

On exposure to air, the surface of magnesium ribbon is oxidized to form a thin film of Magnesium oxide

. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

(ii)Burning in air/oxygen

Experiment

Hold a about 2cm length of Magnesium ribbon on a Bunsen flame. Stop heating when it catches fire/start burning.

Caution: Do not look directly at the flame

Put the products of burning into 100cm3 beaker. Add about 5cm3 of distilled water. Swirl. Test the mixture using litmus papers.

Repeat with Calcium

Observations

-Magnesium burns with a bright blindening flame
-White solid /ash produced
-Solid dissolves in water to form a colourless solution
-Blue litmus paper remain blue
-Red litmus paper turns blue
-colourless gas with pungent smell of urine

Explanation

Magnesium burns in air with a bright blindening flame to form a mixture of Magnesium oxide and Magnesium nitride.

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

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

Magnesium oxide dissolves in water to form magnesium hydroxide.

$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$

Magnesium nitride dissolves in water to form magnesium hydroxide and produce ammonia gas.

$Mg_3N_2(s) + 6H_2O(l) \rightarrow 3Mg(OH)_2(aq) + 2NH_3(g)$

Magnesium hydroxide and ammonia are weakly alkaline with pH 8/9/10/11 and turns red litmus paper blue.

Calcium burns in air with faint orange/red flame to form a mixture of both Calcium oxide and calcium nitride.

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

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

Calcium oxide dissolves in water to form calcium hydroxide.

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

Calcium nitride dissolves in water to form calcium hydroxide and produce ammonia gas.

$Ca_3N_2(s) + 6H_2O(l) \rightarrow 3Ca(OH)_2(aq) + 2NH_3(g)$

Calcium hydroxide is also weakly alkaline solution with pH 8/9/10/11 and turns red litmus paper blue.

(iii)Reaction with water

Experiment

Measure 50 cm3 of distilled water into a beaker.
Scrub/polish with sand paper 1cm length of Magnesium ribbon
Place it in the water. Test the product-mixture with blue and red litmus papers.
Repeat with Calcium metal.

Observations

-Surface of magnesium covered by bubbles of colourless gas.
-Colourless solution formed.
-Effervescence/bubbles/fizzing takes place in Calcium.
-Red litmus paper turns blue.
-Blue litmus paper remains blue.

Explanations

Magnesium slowly reacts with cold water to form Magnesium hydroxide and bubbles of Hydrogen gas that stick on the surface of the ribbon.

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

Calcium moderately reacts with cold water to form Calcium hydroxide and produce a steady stream of Hydrogen gas.

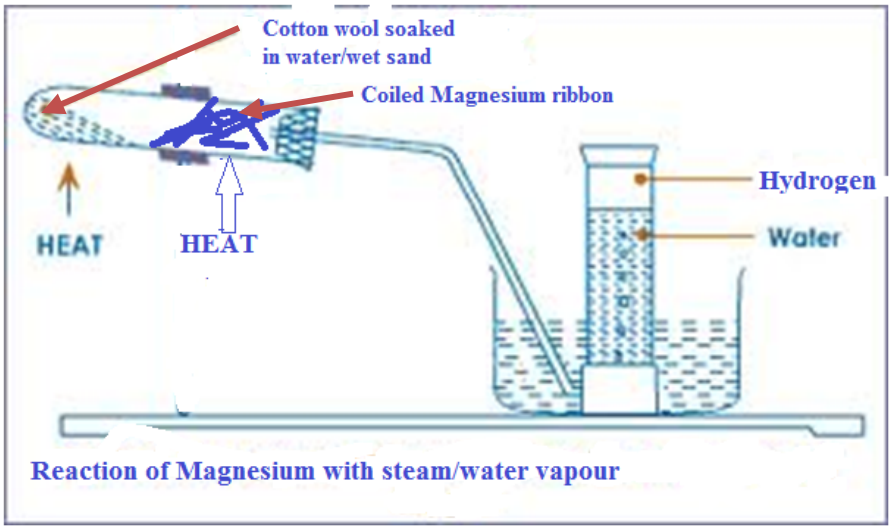
$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

(iv)Reaction with water vapour/steam

Experiment

Put some cotton wool soaked in water/wet sand in a long boiling tube.
Coil a well polished magnesium ribbon into the boiling tube.
Ensure the coil touches the side of the boiling tube. Heat the cotton wool/sand slightly then strongly heat the Magnesium ribbon .

Set up of apparatus



Observations

-Magnesium glows red hot then burns with a blindening flame.
-Magnesium continues to glow/burning even without more heating.
-White solid/residue.
-colourless gas collected over water.

Explanation

On heating wet sand, steam is generated which drives out the air that would otherwise react with /oxidize the ribbon.
Magnesium burns in steam/water vapour generating enough heat that ensures the reaction goes to completion even without further heating. White Magnesium oxide is formed and hydrogen gas is evolved.
To prevent suck back, the delivery tube should be removed from the water before heating is stopped at the end of the experiment.

$Mg(s) + H_2O(l) \rightarrow MgO(s) + H_2(g)$

(v)Reaction with chlorine gas.

Experiment

Lower slowly a burning magnesium ribbon/shavings into a gas jar containing Chlorine gas. Repeat with a hot piece of calcium metal.

Observation

-Magnesium continues to burn in chlorine with a bright blindening flame.
-Calcium continues to burn for a short time.
-White solid formed .
-Pale green colour of chlorine fades.

Explanation

Magnesium continues to burn in chlorine gas forming white magnesium oxide solid.

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

Calcium burns slightly in chlorine gas to form white calcium oxide solid. Calcium oxide formed coat unreacted Calcium stopping further reaction

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

(v)Reaction with dilute acids.

Experiment

Place about 4.0cm3 of 0.1M dilute sulphuric(VI)acid into a test tube. Add about 1.0cm length of magnesium ribbon into the test tube. Cover the mouth of the test tube using a thumb. Release the gas and test the gas using a burning splint.
Repeat with about 4.0cm3 of 0.1M dilute hydrochloric/nitric(V) acid.
Repeat with 0.1g of Calcium in a beaker with all the above acid
Caution: Keep distance when using calcium

Observation

-Effervescence/fizzing/bubbles with dilute sulphuric(VI) and nitric(V) acids
-Little Effervescence/fizzing/bubbles with calcium and dilute sulphuric(VI) acid.
-Colourless gas produced that extinguishes a burning splint with an explosion/ "pop" sound.
-No gas is produced with Nitric(V)acid.
-Colourless solution is formed.

Explanation

Dilute acids react with alkaline earth metals to form a salt and produce hydrogen gas.
Nitric(V)acid is a strong oxidizing agent. It quickly oxidizes the hydrogen produced to water.
Calcium is very reactive with dilute acids and thus a very **small** piece of very **dilute** acid should be used.

Chemical equations

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

$Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2(g)$
(insoluble $CaSO_4$(**s**) coat/cover Ca(s))
$Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$
$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$

$Ba(s) + H_2SO_4(aq) \rightarrow BaSO_4(s) + H_2(g)$
(insoluble $BaSO_4$(**s**) coat/cover Ba(s))
$Ba(s) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + H_2(g)$

$Ba(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2(g)$

The table below shows some compounds of some alkaline earth metals

	Beryllium	Magnesium	Calcium	Barium
Hydroxide	$Be(OH)_2$	$Mg(OH)_2$	$Ca(OH)_2$	$Ba(OH)_2$
Oxide	BeO	MgO	CaO	BaO
Sulphide	-	MgS	CaS	BaS
Chloride	$BeCl_2$	$MgCl_2$	$CaCl_2$	$BaCl_2$
Carbonate	$BeCO_3$	$MgCO_3$	$CaCO_3$	$BaCO_3$
Nitrate(V)	$Be(NO_3)_2$	$Mg(NO_3)_2$	$Ca(NO_3)_2$	$Ba(NO_3)_2$
Sulphate(VI)	$BeSO_4$	$MgSO_4$	$CaSO_4$	$BaSO_4$
Sulphate(IV)	-	-	$CaSO_3$	$BaSO_3$
Hydrogen carbonate	-	$Mg(HCO_3)_2$	$Ca(HCO_3)_2$	-
Hydrogen sulphate(VI)	-	$Mg(HSO_4)_2$	$Ca(HSO_4)_2$	-

Some uses of alkaline earth metals include:

(i)Magnesium hydroxide is a non-toxic/poisonous mild base used as an anti acid medicine to relieve stomach acidity.
(ii)Making **duralumin.** Duralumin is an alloy of Magnesium and aluminium used for making aeroplane bodies because it is light.
(iii) Making plaster of Paris-Calcium sulphate(VI) is used in hospitals to set a fractures bone.
(iii)Making cement-Calcium carbonate is mixed with clay and sand then heated to form cement for construction/building.
(iv)Raise soil pH-Quicklime/calcium oxide is added to acidic soils to neutralize and raise the soil pH in agricultural farms.
(v)As nitrogenous fertilizer-Calcium nitrate(V) is used as an agricultural fertilizer because plants require calcium for proper growth.
(vi)In the blast furnace-Limestone is added to the blast furnace to produce more reducing agent and remove slag in the blast furnace for extraction of Iron.

(c)Group VII elements: Halogens

Group VII elements are called **Halogens.** They are all non metals. They include:

Element	Symbol	Atomic number	Electronicc configuration	Charge of ion	Valency	State at Room Temperature
Fluorine	F	9	2:7	F^-	1	Pale yellow gas
Chlorine	Cl	17	2:8:7	Cl^-	1	Pale green gas
Bromine	Br	35	2:8:18:7	Br^-	1	Red liquid
Iodine	I	53	2:8:18:18:7	I^-	1	Grey Solid
Astatine	At	85	2:8:18:32:18:7	At^-	1	Radioactive

All halogen atoms have seven electrons in the outer energy level. They **acquire/gain one** electron in the outer energy level to be stable. They therefore are therefore **monovalent** .They exist in oxidation state X^-
The number of energy levels increases down the group from Fluorine to Astatine. The more the number of energy levels the bigger/larger the atomic size. e.g.
The atomic size/radius of Chlorine is bigger/larger than that of Fluorine because Chlorine has more/3 energy levels than Fluorine (2 energy levels).

Atomic radius and ionic radius of Halogens increase down the group as the number of energy levels increases.

The atomic radius of Halogens is smaller than the ionic radius. This is because they react by gaining/acquiring extra one electron in the outer energy level. The effective nuclear attraction on the more/extra electrons decreases. The incoming extra electron is also repelled causing the outer energy level to expand to reduce the repulsion and accommodate more electrons.

Table showing the atomic and ionic radius of four Halogens

Element	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

The atomic radius of Chlorine is 0.099nM .The ionic radius of Cl^- is 0.181nM. This is because Chlorine atom/molecule reacts by gaining/acquiring extra one electrons. The more/extra electrons/energy level experience less effective nuclear attraction /pull towards the nucleus .The outer enegy level expand/increase to reduce the repulsion of the existing and incoming gained /acquired electrons.
Electronegativity
The ease of gaining/acquiring extra electrons is called electronegativity. All halogens are electronegative. Electronegativity decreases as atomic radius

increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius.
The outer electrons experience less nuclear attraction and thus ease of gaining/acquiring extra electrons decrease.
It is measured using Pauling's scale.
Where Fluorine with Pauling scale 4.0 is the most electronegative element and thus the highest tendency to acquire/gain extra electron.

Table showing the electronegativity of the halogens.

Halogen	F	Cl	Br	I	At
Electronegativity (Pauling scale)	4.0	3.0	2.8	2.5	2.2

The electronegativity of the halogens decrease down the group from fluorine to Astatine. This is because atomic radius increases down the group and thus decrease electron – attracting power down the group from fluorine to astatine.

Fluorine is the most electronegative element in the periodic table because it has the small atomic radius.

Electron affinity

The minimum amount of energy required to gain/acquire an extra electron by an atom of element in its gaseous state is called **1st electron affinity**. The SI unit of electron affinity is **kilojoules per mole/kJmole^{-1} .** Electron affinity depend on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer energy level electrons and thus the lower the electron affinity. For halogens the 1st electron affinity decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease. Due to its small size/atomic radius Fluorine shows exceptionally low electron affinity. This is because a lot of energy is required to overcome the high repulsion of the existing and incoming electrons.

Table showing the election affinity of halogens for the process

$X + e \rightarrow X^-$

Halogen	F	Cl	Br	I
Electron affinity kJmole^{-1}	-333	-364	-342	-295

The higher the electron affinity the more stable theion.i.e
Cl^- is a more stable ion than Br^- because it has a more negative / exothermic electron affinity than Br^-

Electron affinity is different from:
(i) Ionization energy.

Ionization energy is the energy required to lose/donate an electron in an atom of an element in its gaseous state while electron affinity is the energy required to gain/acquire extra electron by an atom of an element in its gaseous state.

(ii) Electronegativity.

-Electron affinity is the energy required to gain an electron in an atom of an element in gaseous state. It involves the process:

$$X(g) + e \rightarrow X^-(g)$$

Electronegativity is the ease/tendency of gaining/ acquiring electrons by an element during chemical reactions.

It does not involve use of energy but theoretical arbitrary Pauling' scale of measurements.

Physical properties

State at room temperature

Fluorine and Chlorine are gases, Bromine is a liquid and Iodine is a solid. Astatine is radioactive .

All halogens exist as **diatomic** molecules bonded by strong covalent bond. Each molecule is joined to the other by weak intermolecular forces/ Van-der-waals forces.

Melting/Boiling point

The strength of intermolecular/Van-der-waals forces of attraction increase with increase in molecular size/atomic radius.

Iodine has therefore the largest atomic radius and thus strongest intermolecular forces to make it a solid.

Iodine sublimes when heated to form (**caution:** highly toxic/poisonous) purple vapour.

This is because Iodine molecules are held together by weak van-der-waals/intermolecular forces which require little heat energy to break.

Electrical conductivity

All Halogens are poor conductors of electricity because they have no free delocalized electrons.

Solubility in polar and non-polar solvents

All halogens are soluble in water(polar solvent).

When a boiling tube containing either chlorine gas or bromine vapour is separately inverted in a beaker containing distilled water and tetrachloromethane (non-polar solvent), the level of solution in boiling tube rises in both water and tetrachloromethane.

This is because halogen are soluble in both polar and non-polar solvents. Solubility of halogens in water/polar solvents decrease down the group. Solubility of halogens in non-polar solvent increase down the group.

The level of water in chlorine is higher than in bromine and the level of

tetrachloromethane in chlorine is lower than in bromine.
Caution: Tetrachloromethane , Bromine vapour and Chlorine gas are all **highly** toxic/poisonous.

Table showing the physical properties of Halogens

Halogen	Formula of molecule	Electrical conductivity	Solubility in water	Melting point(^oC)	Boiling point(^oC)
Fluorine	F_2	Poor	Insoluble/soluble in tetrachloromethane	-238	-188
Chlorine	Cl_2	Poor	Insoluble/soluble in tetrachloromethane	-101	-35
Bromine	Br_2	Poor	Insoluble/soluble in tetrachloromethane	7	59
Iodine	I_2	Poor	Insoluble/soluble in tetrachloromethane	114	sublimes

Chemical properties

(i)Displacement

Experiment
Place separately in test tubes about $5cm^3$ of sodium chloride, Sodium bromide and Sodium iodide solutions.
Add 5 drops of chlorine water to each test tube:
Repeat with 5 drops of bromine water instead of chlorine water
Observation
Using Chlorine water
-Yellow colour of chlorine water fades in all test tubes except with sodium chloride.
-Coloured Solution formed.
Using Bromine water
Yellow colour of bromine water fades in test tubes containing sodium iodide.
-Coloured Solution formed.
Explanation
The halogens displace each other from their solution. The more electronegative displace the less electronegative from their solution.
Chlorine is more electronegative than bromine and iodine.
On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.
Bromine is more electronegative than iodide but less 6than chlorine.
On adding Bromine water, iodine is displaced from its solution but not chlorine.

Table showing the displacement of the halogens

(V) means there is displacement (x) means there is no displacement

Halogen ion in solution / Halogen	F^-	Cl^-	Br^-	I^-
F_2	X			
Cl_2	X	X		
Br_2	X	X	X	
I_2	X	X	X	X

Chemical /ionic equations

With Fluorine

$F_2(g) + 2NaCl^-(aq) \rightarrow 2NaF(aq) + Cl_2(aq)$

$F_2(g) + 2Cl^-(aq) \rightarrow 2F^-(aq) + Cl_2(aq)$

$F_2(g) + 2NaBr^-(aq) \rightarrow 2NaF(aq) + Br_2(aq)$

$F_2(g) + 2Br^-(aq) \rightarrow 2F^-(aq) + Br_2(aq)$

$F_2(g) + 2NaI^-(aq) \rightarrow 2NaF(aq) + I_2(aq)$

$F_2(g) + 2I^-(aq) \rightarrow 2F^-(aq) + I_2(aq)$

With chlorine

$Cl_2(g) + 2NaCl^-(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$

$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$

$Cl_2(g) + 2NaI^-(aq) \rightarrow 2NaCl(aq) + I_2(aq)$

$Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

With Bromine

$Br_2(g) + 2NaI^-(aq) \rightarrow 2NaBr(aq) + I_2(aq)$

$Br_2(g) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$

Uses of halogens

(i) Florine – manufacture of P.T.F.E (Poly tetra fluoroethene) synthetic fiber.

- Reduce tooth decay when added in small amounts/quantities in tooth paste.

NB – large small quantities of fluorine /fluoride ions in water cause browning of teeth/flourosis.

- Hydrogen fluoride is used to engrave words /pictures in glass.

(ii) Bromine - Silver bromide is used to make light sensitive photographic paper/films.

(iii) Iodide – Iodine dissolved in alcohol is used as medicine to kill bacteria in skin cuts. It is called tincture of iodine.

The table below to show some compounds of halogens.

Element / Halogen	H	Na	Mg	Al	Si	C	P
F	**HF**	**NaF**	**MgH_2**	**AlF_3**	**SiF_4**	**CF_4**	**PF_3**
Cl	**HCl**	**NaCl**	**MgCl**	**$AlCl_3$**	**$SiCl_3$**	**CCl_4**	**PCl_3**
Br	**HBr**	**NaBr**	**$MgBr_2$**	**$AlBr_3$**	**$SiBr_4$**	**CBr_4**	**PBr_3**
I	**HI**	**NaI**	**MgI_2**	**AlI_3**	**SiI_4**	**Cl_2**	**PBr_3**

(i) Below is the table showing the bond energy of four halogens.

Bond	Bond energy k J $mole^{-1}$
Cl-Cl	242
Br-Br	193
I-I	151

I. What do you understand by the term “bond energy”
Bond energy is the energy required to break/ form one mole of chemical bond

II. Explain the trend in bond Energy of the halogens above:
-Decrease down the group from chlorine to Iodine
-Atomic radius increase down the group decreasing the energy required to break the covalent bonds between the larger atom with reduced effective nuclear @ charge an outer energy level that take part in bonding.

(c)Group VIII elements: Noble gases

Group VIII elements are called **Noble gases.** They are all non metals. Noble gases occupy about 1.0% of the atmosphere as colourless gaseous mixture. Argon is the most abundant with 0.9%.
They exists as **monatomic** molecules with very weak van-der-waals /intermolecular forces holding the molecules.
They include:

Element	Symbol	Atomic number	Electron structure	State at room temperature
Helium	He	2	2:	Colourless gas

Neon	Ne	10	2:8	Colourless gas
Argon	Ar	18	2:8:8	Colourless gas
Krypton	Kr	36	2:8:18:8	Colourless gas
Xenon	Xe	54	2:8:18:18:8	Colourless gas
Radon	Rn	86	2:8:18:32:18:8	Radioctive

All noble gas atoms have a stable duplet(two electrons in the 1st energy level) or octet(eight electrons in other outer energy level)in the outer energy level. They therefore **do not** acquire/gain extra electron in the outer energy level or donate/lose. They therefore are therefore **zerovalent** .

The number of energy levels increases down the group from Helium to Randon. The more the number of energy levels the bigger/larger the atomic size/radius. e.g. The atomic size/radius of Argon is bigger/larger than that of Neon because Argon has more/3 energy levels than Neon (2 energy levels).

Atomic radius noble gases increase down the group as the number of energy levels increases.
The effective nuclear attraction on the outer electrons thus decrease down the group.

The noble gases are generally unreactive because the outer energy level has the stable octet/duplet. The stable octet/duplet in noble gas atoms lead to a comparatively very high 1st ionization energy. This is because losing /donating an electron from the stable atom require a lot of energy to lose/donate and make it unstable.
As atomic radius increase down the group and the 1st ionization energy decrease, very electronegative elements like Oxygen and Fluorine are able to react and bond with lower members of the noble gases.e.g
Xenon reacts with Fluorine to form a covalent compound XeF_6.This is because the outer electrons/energy level if Xenon is far from the nucleus and thus experience less effective nuclear attraction.

Noble gases have low melting and boiling points. This is because they exist as monatomic molecules joined by very weak intermolecular/van-der-waals forces that require very little energy to weaken and form liquid and break to form a gas. The intermolecular/van-der-waals forces increase down the group as the atomic radius/size increase from Helium to Radon. The melting and boiling points thus increase also down the group.
Noble gases are insoluble in water and are poor conductors of electricity.

Element	Formula of molecule	Electrical conductivity	Solubility in water	Atomic radius(nM)	1st ionization energy	Melting point(^{0}C)	Boiling point(^{0}C)
Helium	He	Poor	Insoluble	0.128	2372	-270	-269

Neon	Ne	Poor	Insoluble	0.160	2080	-249	-246
Argon	Ar	Poor	Insoluble	0.192	1520	-189	-186
Krypton	Kr	Poor	Insoluble	0.197	1350	-157	-152
Xenon	Xe	Poor	Insoluble	0.217	1170	-112	-108
Radon	Rn	Poor	Insoluble	0.221	1134	-104	-93

Uses of noble gases

Argon is used in light bulbs to provide an inert environment to prevent oxidation of the bulb filament

Argon is used in arch welding as an insulator.

Neon is used in street and advertisement light

Helium is mixed with Oxygen during deep sea diving and mountaineering.

Helium is used in weather balloon for meteorological research instead of Hydrogen because it is unreactive/inert.Hydrogen when impure can ignite with an explosion.

Helium is used in making thermometers for measuring very low temperatures.

TOPICAL QUESTIONS ON THE TOPIC BASED ON PAST KCSE PAPERS

1. The table below gives the atomic numbers of elements W X Y and Z. The letters do not represent the actual symbols of the elements.

Element	W	X	Y	Z
Atomic numbers	9	10	11	12

a) Which one of the elements is less reactive? Explain. (2mks)

b) i) Which two elements would react most vigorously with each other

ii) Give the formula of the compound formed when elements in b (i) above react (1mk)

2. The table below gives the energy required to remove the outer most electrons from same group

Elements	I	II	III	IV
Energy kj /Mole	494	418	519	376

Arrange the electrons in the order of their reactivity starting with the most reactive. (2mks)

3. The information below relates to elements s, T, U, and x. The letters do not represent the actual symbols of the elements.

i) "T" displaces "X" from aqueous solution containing ions of "X"

ii) Hydrogen gases reduces heated oxide of "s" but does not reduce the heated oxide of "X"\

iii) "U" liberates hydrogen gas from cold water but "T" does not

a) Write an equation for the reaction between "T" and ions of "X" both T and X are in the group II of the periodic table (1mk)

b) Arrange the elements in order of their increasing reactivity (1mk)

4. The electronic structures for elements represented by letters A, B, C, and D are:-

A= 2, 8, 6 B= 2, 8, 2 C= 2, 8, 1 D= 2, 8, 8

a) Select the element which forms

i) Double charged cation (1mk)

ii) A soluble carbonate (1mk)

b) Which element has the smallest atomic radius (1mk)

5. The information in the table below relates to elements in the same group of the periodic table. Study it and answer the questions that follows:-

Elements	Atomic size (mm)
G1	0.19
G2	0.23
G3	0.15

Which element has highest ionization energy? Give a reason. (3mks)

6. The oxides of elements “A” and “B” have the properties shown in the table below. The letters do not represent actual symbols of the elements.

A	B
A gas at room temperature	Solid normal temperature
Dissolves in water to form acidic solution	Dissolves in water to form alkaline solution

Give one example of element “A” and “B” (2mks)

7. An oxide of F has the formula F_2O_5
   a) Determine the oxidation state of “F” (1mk)
   b) In which group of the periodic table is element “F” (1mk)
8. Yellow phosphorus reacts with chlorine gas to form a yellow liquid. The liquid fumes when exposed to air. Explain these observations. (2mks)
9. Explain why the reactivity of group (VII) elements decreases down the group. (3mks)
10. The atomic numbers of element “C” and “D” are 19 and 9 respectively. State and explain the electro conductivity of compound CD in:-
    a) Solid state (1 ½ mark)
    b) Aqueous state (1 ½ mark)
11. a) Explain why the metals magnesium and aluminium are good conductors of electricity. (1mk)
    b) Other than cost, give two reasons why aluminium is used for making electric cables while magnesium is not. (2mks)
12. The table below gives information on four elements represented by letters K, L, M and N. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

Elements	Electron arrangement	Atomic radius (nm)	Ionic radius
K	2,8,2	0.136	0.065
L	2,8,7	0.099	0.181
M	2,8,8,1	0.099	0.181
N	2,8,8,2	0.174	0.099

    a) Which two elements have similar chemical properties? Explain (2mks)
    b) What is the most likely formula of the oxide of “L” (1mk)
    c) Which element is a non-metal? Explain (2mks)
    d) Which one of the elements is the strongest reducing agent? Explain (2mks)
    e) Explain why the ionic radius of “N” is less than that of “M” (2mks)
    f) Explain why the ionic radius of “L” is larger than its atomic radius. (2mks)

13. Study the information given in the table below and answer the questions that follow. The letters do not represent the actual symbols of elements.

Elements	Atomic numbers	Boiling point
S	3	1603
T	13	2743
U	16	718
V	18	87
W	19	1047

a) Select the element which belong to the same
   i) Group (1mk)
   ii) Period (1mk)

b) Which element
   i) is in gaseous state at room temperature? Explain (2mks)
   Take room temperature to be 298K
   ii) Does not form oxides (1mk)

c) Write the:-
   i) Formula of the nitrate of element T (1mk)
   ii) Equation for the reaction between element "S" and "U" (1mk)

d) What type of bond would exist in the compound formed when element "U" and "T" react? Give a reason for your answer (2mks)

e) The aqueous sulphate of element "w" was electrolyzed using inert electrodes. Name the products formed at the
   i) Cathode (1mk)
   ii) Anode (1mk)

14. The table below shows some properties of chlorine, bromine and iodine.

Elements	Formulae	Colour and state at room temperature	Solubility in water
Chlorine	Cl_2	(i)…….	Soluble
Bromine	Br_2	Brown liquid	(ii)……
Iodine	I_2	(iii) ……	Slightly soluble

a) Complete the table below by giving the missing information in (i) (ii) (3mks)

b) Chloride is prepared by reacting concentrated hydrochloric acid with Manganese (IV) oxide.
   i) Write the equation for the reaction between concentrated hydrochloric acid and manganese (IV) oxide.
   ii) What is the role of manganes (IV) oxide in this reaction (1mk)

c) i) Iron (ii) chloride reacts with chlorine gas to form substance "E". Identify substance "E" (1mk)
   ii) During the reaction in c (i) above, 6.30g of iron (II) chloride were converted to 8.06g of substance "E". Calculate the volume of chlorine gas used. (Cl=35.5) molar gas at room temperature = 24000 cm^3 (Fe= 56)

(3mks)

d) Draw and name the structure of the compound formed when excess chlorine gas is reacted with ethane gas. (2mks)

Structure…….

Name …….

15. The grid below represents part of the periodic table. Study it and answer the questions that follows:- The letter given do not represent the actual symbols of the elements.

						A		
	B		C		D		E	
F	G							
							H	

i) Select the element that can form an ion with a change of-2. Explain your answer. (2mks)

ii) What type of structure would the oxide of C have? Explain your answer. (2mks)

iii) How does reaction of H compare with that of E? (2mks)

iv) 1.3g of "B" react completely when heated with 1.21 litres of $Cl_{2\,(g)}$ at STP. (1 mole of gas of STP occupies 22.4 litres)

I) Write a balanced equation for the reaction between B and Cl_2(1mk)

Ii) Determine the relative atomic mass of B. (2mks)

v) Explain how you would expect the following to compare.

a) Atomic radii of "F" and "G" (1mk)

b) The pH values of aqueous solution of oxides of B and D. (2mks)

vi) The table below shows some physical properties of some substances. Use the information in the table to answer the questions that follow:-

				Electrical conductivity
Substances	Melting	Boiling point ^0C	Solid	Solid
U	1083	2595	Good	Good
V	801	1413	Poor	Good
W	5.5	80.1	Poor	Poor
X	-114.8	-84.9	Poor	Poor
Y	3550	4827	Poor	poor

i) Which substance is likely to be (1mk)

(I) A metal (1mk)

(II) Liquid at room temperature (1mk)

ii) Which substance is likely to have the following structures?

(I) Simple molecular (1mk)

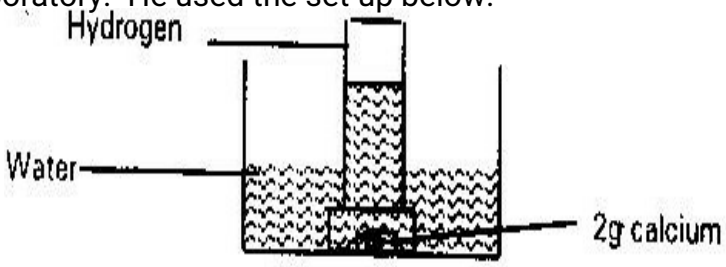
(II) Giant atomic (1mk)

16. Lithium, sodium and potassium belong to the same group of the periodic table

i) Arrange the elements in the order of increasing ionization energy. (1mk)

ii) Explain the trend in 2(i) above (2mks)

17. When heated in a current of Nitrogen gas, magnesium reacts to form a compound magnesium nitride, Mg_3N_2
    a) Calculate a volume of Nitrogen at s.t.p required to react with 8g of magnesium (Mg= 24) molar gas volume at s.t.p= 22.4 dm^3) (3mks)
    b) Magnisium Nitrite reacts with water to form magnesium hydroxide and ammonia. Calculate the volume of ammonia produced at S.T.P, if all magnesium nitride formed reacts completely with water. (3mks)
18. A student at Loreto Secondary school used 2g of calcium to prepare hydrogen gas in the laboratory. He used the set up below.



    a) Write a chemical equation for the reaction that produced hydrogen (1mk)
    b) Calculate the volume of hydrogen produced at room temperature (molar gas volume= 24,000cm^3) (2mks)
    c) Explain why the same method cannot be used to prepare hydrogen using sodium in the laboratory (2mks)
    d) Explain why the same method cannot be used to prepare hydrogen using sodium in laboratory (2mks)
    e) Calculate the mass of the products formed if all the hydrogen produced in this experiment was burnt in excess air. (3mks)
    f) Explain how calcium is able to conduct electricity (2mks)
19. The table below gives atomic and mass numbers of some elements represented by letters "T" to "Y". The letters are not actual symbols of elements. Use it to answer questions that follows:-

Elements	T	U	V	W	X	Y
Atomic numbers	1	18	1	19	20	17
Mass numbers	2	39	1	39	40	35

    a) Which element has the lowest ionization energy? (2mks)
    b) Element "V" is uniquely positioned in the periodic table. It has a tendency of forming compounds by either gaining or sharing electrons. Give the formula of a compound of "V" that is formed when V gain an electron. (1mk)
20. When magnesium metal burn metal burn in air. It reacts with both oxygen and Nitrogen gases giving a white ash- like substances. Write two equations for the two reactions that take place. (2mks)
21. Chlorine and iodine are elements in the same group in the periodic table. Chlorine gas is yellow while iodine solution is brown.
    a) What observations would be made if chlorine gas is bubbled through aqueous sodium iodide? Explain using an ionic equation. (1mk)

b) Under certain conditions chlorine and iodine react to give iodine trichloride ($LCl_{3\,(s)}$). What type of bonding would you expect to exist in iodine trichloride? Explain. (1mk)

22. It is not appropriate to refer to group VIII elements as "inert gases" Explain giving an example. (2mks)
23. What observations will you make when chlorine gas is bubbled through
    i) Potassium bromide (1mk)
    ii) Potassium chloride (1mk)
    iii) Explain these observations (3mks)
24. Explain why the reactivity of group (VIII) elements decreases down the group. (3mks)

ANSWERS

1. a) X: it has a stable electronic arrangement i.e Octet structure.
   b) i) "W" and "Y"
   ii) YW
2. IV, II, I, III
3. a) $T_{(s)} + X^{+}_{(aq)} \rightarrow T^{2+}_{(aq)} + X_{(s)}$
   b) S, X, T, U
4. a) i) B ii) C b) D
5. G 3, it has the smallest atomic size, therefore outermost electrons are strongly attracted by the nucleus. A lot of energy is required to remove the outermost electron.
6. Element A= Sulphur, carbon, nitrogen
   Element B = Sodium, potassium, lithium
7. a) F_2O_5 = O: 2F + - 10 = 0 : 2F = 0+10
   $F = \frac{+10}{2} = +5$
   b) Group V
8. The yellow liquid is pcl_3. It is hydrolised in air to form Hcl which fumes since it absorbs water vapour from the atmosphere.
9. Group (VII) elements react by gaining the electrons Flourine has the smallest atomic radius in this group hence it attract electrons very strongly hence it gain electrons very easily making it to be more reactive. Ease of electron gain decreases down the group.
10. a) Solid CD does not conduct electricity since the ions are not free to move. The ions are held together by electrostatic force of attraction.
    b) Aqueous CD is a strong electrolyte since the ions are free to move.
11. a) The outermost electrons in mg and Al are delocalized and free to move hence allow the flow of electric current.
    b) Alluminium forms a protective coating and prevents further corrosion.
12. a) "K" and "N" they are in the same group or same number of valency electron/or they loose two electrons.
    b) L_2O
    c) "L" it has 7 electrons in its outermost energy level/ react by gaining one electron. Its ionic radius is bigger than atomic radius.
    d) M; It has highest tendency to loose electrons.
    e) The ions of "N" have many protons in its nucleus compared to M. The protons in N nucleus pulls the electrons very close to its nucleus.
    f) "L" gains electrons to form L^- ion, the added electron increases the repulsion/screening effect between electrons in the adjacent energy levels.
13. a) i) "S" and "W"
    b) i) "V" it is the only element whose boiling point is below 298 kj at room temperature.
    ii) V has stable electron arrangement
    c) i) $T(NO_3)_3$
    ii) $2S_{(s)} + U_{(s)} \rightarrow S_2U_{(s)}$

d) Ionic or electrovalent bond "T" is a metal while "U" is a non metal, therefore T loses electrons to "U"

e) i) Cathode- Hydrogen gas

ii) Anode – Oxygen gas.

14. a) i) Greenish yellow gas

ii) Slightly soluble

iii) Grey/ black solid.

b) (i) $4HCl_{(aq)} + MnO_{2(s)} \rightarrow MnCl_{2(aq)} + 2HO_{(l)} + Cl_{2(g)}$

ii) Oxidizing agent. It oxidizes the chloride ions to chlorine gas.

c) i) Iron (iii) Chloride

ii) Mass of chlorine used

= 8.06 – 6.30 = 1.76g

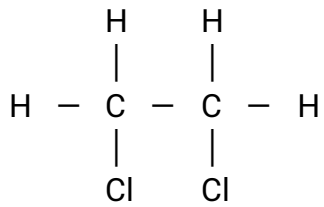
RM M of Cl_2 = 71

Moles of cl_2 = $\frac{1.77}{71}$ = 0.0248 moles

Volume of chlorine

= 0.0248 x 24000 = 595.2 cm^3

d)



1, 2 – dichloroethane

15. i) "A" It is in group (VI) and gaining two electrons .

ii) Giant ionic structure: C_2O_3 is an ionic compound. This is a very strong force of attraction (electrostatic force) between the ions.

iii) "E" is more reactive than H. "E" has a small atomic radii and gains electrons very easily compared to H.

iv) (I) $B_{(s)} + Cl_{2(g)} \rightarrow BCl_{2(s)}$

(II) Moles of Cl_2 = $\frac{1.21}{22.4}$ = 0.054 moles

Moles of B = 0.054 moles

RAM of = $\frac{1.3}{0.054}$ = 24

v) "G" has a small atomic radius compared to F. G has many protons and hence attracts electrons very easily to its nuleus.

b) i) The oxide of B is alkaline in mature with a PH greater than (8.0).

B is a metal and forms basic oxide. D is a non metal and forms acidic oxide with a PH less than 5.0

ii) i) I U

II W
ii) I X
II Y

16. i) Potassium – sodium – Lithium
ii) Lithium has a small atomic radius compared to the others. The outermost electrons are attracted very strongly by the nucleus charges.
A lot of energy is required to pull out the outer most electrons. Atomic radius decreases from potassium to lithium.

17. a) $3\ Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_2$
Moles of magnesium = $\frac{8}{24}$ = 0.333 moles
Moles of N_2 = $\frac{0.333}{3}$ = 0.111 moles
Volume of N_2 = 0.111 x 22.4 = 2.488dm^3
b) $Mg_3N_{2(s)} + 6H_2O_{(l)} \rightarrow 3mg\ (OH)_{2(aq)} + 2MH_{3(g)}$
Moles of Mg_3N_2 =0.111 moles
Moles of NH_3 = 0.111 x 2 = 0.222 moles.
Volume of NH_3 = 0.222 x 22.4 = 4.97dm^3

18. a) $Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca\ (OH)_{2\ (aq)} + H_{2(g)}$
b) Moles of Ca = $\frac{2}{40}$ = 0.05 moles
Moles of H_2 ≡0.05 moles
Volume of H_2 =0.05 x 24000 =1,200 cm^3
c) $Ca\ (OH)_2$ is slightly soluble in water
d) Sodium reacts with water very vigorously. Reaction may end being explosive since sodium is very reactive.
e) $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O$
Moles of H_2 = 0.05 moles
Moles of H_2O ≡ 0.05 moles
RMM of H_2O =18
Mass of H_2 = 0.005 x 18 = 0.9g
f) Calcium is a metal and the outer most electrons are delocalized/ free to move.

19. a) "W" it has the largest atomic radius. The outermost electrons are loosely held by the nucleus. Less energy is required to remove this electron.
b) V+ X →V + X
V + Na →NaV

20. $2Mg_{(s)} + N_{2(g)} \rightarrow 2MgO_{(s)}$
$3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$

21. a) Grey precipitate of iodine will be observed. Chlorine is more reactive than iodine and it displaces it from its solution of sodium iodide.
$Cl_{2(s)} + 2I_{(aq)} \rightarrow I_{2(s)} + 2Cl_{(aq)}$
(b) Covalent bond both chlorine and iodine are non metals and react by

sharing electrons.

22. Elements in group (VIII) which have a big atomic radius can react under special condition by losing electrons e.g xenon and fluorine- can react to give xenon hexafluoride
23. i) Reddish brown liwuid
    ii) No change
    iii) Chlorine is more reactive than bromine and can displace it from its salts solution, but chlorine can not displace it self.
24. The group (VIII) element reacts by gaining electrons. The atomic radius decreases down the group. Atoms with small ionic radius gains electrons very easily. Hence gain electrons (electro negativity) decreases with an increase in atomic radius.