Kenya Certificate of Secondary Education

TOP NOTCH CHEMISTRY

Paragon of excellence

Students' Form two Notebook

FIFTH EDITION

© THADDEUS M. MBALUKA 2021

ALLIANCE HIGH SCHOOL

This book belongs to:

Name ………………………………………ADM………

School………………………Class……………………

*"The **chemists are a strange class of mortals**, impelled by an almost insane impulse to seek their pleasures amid smoke and vapour, soot and flame, poisons and poverty; yet among all these evils I seem to live so sweetly that may I die if I were to change places with the Persian king."*
Anonymous

PUBLISHED BY

TOPNOTCH EDUCREATIONS

P. O. Box 9373-00200

NAIROBI Kenya.

topnotchchemistry@gmail.com

+254728951320

+254763951320

© Thaddeus M. Mbaluka

All rights reserved. No part of this publication may be reproduced, stored in a retrieval system or transmitted in any form or by any means electronic, mechanical, photocopying, recording or otherwise without the prior written permission of the publisher.

First published 2017

ISBN: 978-9966-099-87-7

Design and Layout by Thaddeus M. Mbaluka

Cover design by Anderson Mbaga & Thaddeus M. Mbaluka

PRINTED BY
MWANAABAKARI PRINTING PRESS, NAIROBI

0728-951320/0763951320

Forward

One of the major objectives of the 8-4-4 system of education is the acquisition of relevant skills, knowledge and attitudes that will enable the learner to face the challenges of higher education and to be self-reliant after completion of school. This need has necessitated the writing of this hand book.

This handbook has especially captured the needs of students as it has integrated revision exercises after every sub topic hence making it student centered. The authors' alignment of notes cum revision exercises has accommodated the needs of learners hence reaching all students.

This book has meticulously designed and customized to meet the needs of the student. After going through it I can authoritatively confirm it contains well researched notes after thorough brain work and active-thought process. The author has really studied his niche, seen the gaps, and came up with an excellent product which bridge the gaps in chemistry and demystify the subject. The author has also pointed out some common errors where students lose marks during marking.

Lastly the heart of chemistry as a subject is in form two, and this book will ensure the student comprehends the concept through exposure to many practice questions in this book. The book is also very useful for teachers who wish to expose students to exam type of questions; I highly recommend it for students and teachers.

ALEX BOSIRE MOMANYI

HEAD OF CHEMISTRY

ALLIANCE HIGH SCHOOL

ACKNOWLEDGEMENT

This book would not have been written without the support of some individuals who gave me the insights, guidance and assisted me in articulating my ideas and blending them to come up with this piece of work. First, I thank the Almighty God for the many blessings, secondly I thank my wife Christine Mbaika Makau who worked tirelessly with me and typed most of the work and drew the diagrams in this book, I would also like to thank the entire fraternity of Alliance high school and all the students for their input in asking questions and encouraging me in compiling this book.

Special thank you to my friend Richard Awino of St Teresa for your irreplaceable support and inspiration in writing this book. Mr Anderson Mbaga who helped in designing the cover of this book. Special gratitude goes to my friend John Watuka, the chief editor of this book from A.I.C Matetani who worked with me in editing book and for his input. A big thank you To Nicholas Mbeva and Madam Rosemary Musyoki for proof reading this book and making the necessary corrections. Madam Mutheu Masave Bura Girls' for your selfless input in reading through the book and making necessary corrections. My students : Sarah Mbithe ,Jackline Mutende ,Benson Ndambuki ,Richard musyoka and Patrick Musyoki who helped me in reading through the book and making necessary corrections. Special thank you chief editor **DENNIS WEKESA NAKITARE** (alliance high school) for guiding our team of editors and to Madam Rose M. Kongo , st mary's Ngongoni for your irreplaceable guidance on examination techniques in compiling this book.

Lastly I appreciate the support I have got from teachers who have given me a lot of encouragement and appreciated my work Judith wambua ,Kyeni girls ,Sammuel Chege Tala Girls, loise Wangui (Othaya girls) ,John Kamuru and Alex Momanyi of Alliance boys,Stanley Ihinga (ndiani girls secondary) Justus Nzyuko (Kunikila) I appreciate you all.

THADDAEUS MAKAU MBALUKA 0728951320 /0763951320

ALLIANCE HIGH SCHOOL

©2018

TABLE OF CONTENTS

TOPIC 1: PERIODIC, TABLES, ATOMIC STRUCTURES AND CHEMICAL FAMILIES

Specific Objectives

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

a) name and write the chemical symbols of the first twenty elements of the periodic table

b) describe the structure of the atom and write the electron arrangement of the first twenty elements of the periodic table

c) explain the electron arrangement of the atom in terms of cnergy levels

d) define atomic number, mass number, isotopes and relative atomic mass

e) calculate the relative atomic masses from isotopic composition

f) explain the position of an element in the periodic table in terms of the electron arrangement

g) define valency and oxidation number of an element

h) predict the type of ion formed from a given electron arrangement of an atom

i) predict the valencies and oxidation numbers from position of elements in the periodic table

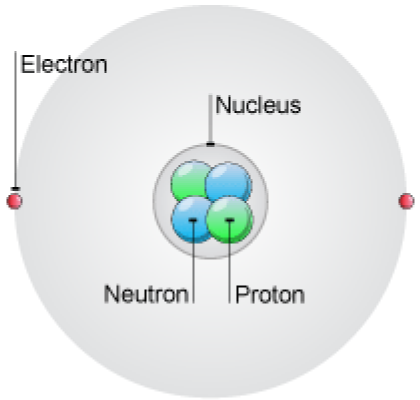
j) derive the formulae of some simple compounds from valences of elements and radicals

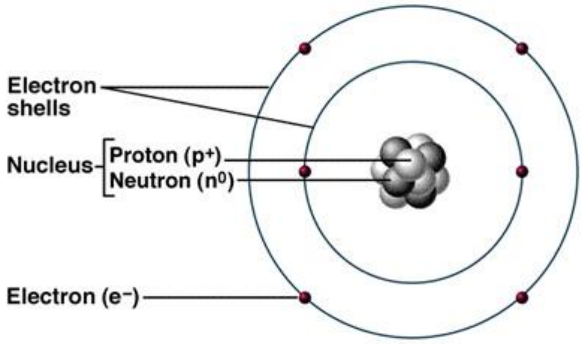
k) Write simple balanced chemical equations.

ATOMIC STRUCTURE

An atom is the smallest particle of an element that can take part in a chemical reaction. The atom is made up of two regions the central part called the **nucleus** and a region surrounding the nucleus which forms the **electron cloud.** It has three subatomic particles: **protons, neutrons** and **electrons**.The nucleus houses the proton and the neutrons while the electron cloud contains electrons. Sub – atomic particles are smaller particles that constitute an atom.

Properties of sub-atomic particles





Summary of subatomic particles

Particle	Relative mass	Symbol	Charge
Proton	1	**p**	+1
Electron	$\frac{1}{1840}$	**e**	–1
Neutron	1	**n**	0

(i) Protons

1. The proton is positively charged.
2. Is found in the centre of an atom called **nucleus.**
3. It has a relative mass 1.
4. The number of protons in a atom of an element is its **atomic number**.

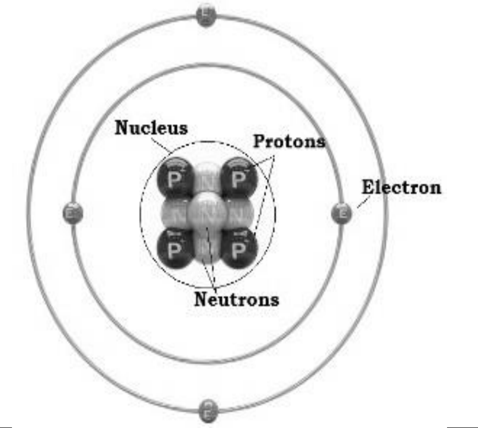
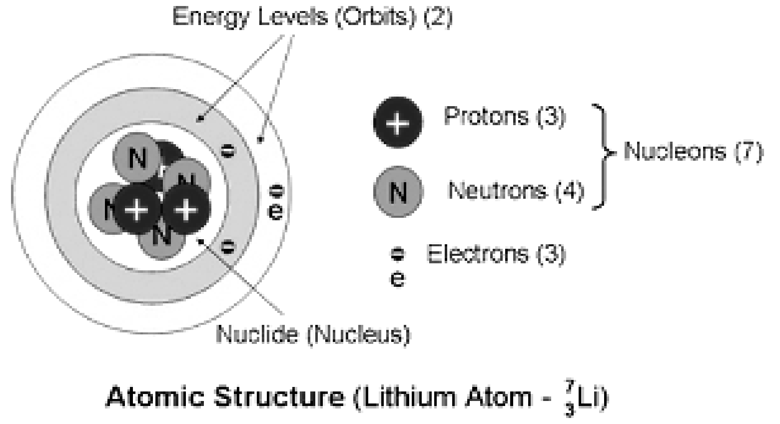
(ii) Electrons

1. The electron is negatively charged.
2. Is found in fixed regions surrounding the centre of an atom called energy levels/shells.
3. It has a relative mass $^{1}/_{1840}$.

(iii) **Neutrons**

1. The neutron is neither positively or negatively charged thus neutral.
2. Like protons, it is found in the centre of an atom called nucleus.
3. It has a relative mass 1.

Diagram showing the relative positions of protons, electrons and neutrons in an atom of helium and lithium

Helium	Lithium
	

NB: In an atom, the number of protons equals to the number of electrons and it's why an atom is electrically neutral.

Atomic number and mass number

Atomic number is the number of protons in an atom.
Mass number is the sum of protons and neutrons in an atom.

NB: In any atom,

- ◦ Atomic number $=$ number of protons $=$ number of electrons
- ◦ Number of Protons $=$ Number of electrons (that is why an atom is electrically neutral)
- ◦ Mass number $=$ protons + neutrons
- ◦ Number of neutrons $=$ mass no – number of protons

When getting the mass of an atom, the mass of electrons is ignored because it is **infinitely smaller** *(very very very small)/neglible.* E.g.

a) ***Magnesium (Mg)*** *has 12 p and 12 n: Its mass number= 12+12= 24*

b) ***Potassium (K)*** *has 20 p and 19 n: Its mass number= 20+19= 39*

When writing the chemical symbol of an element the mass number is written as a **superscript** on the left while the atomic number is written as a **sub script** on the left.

Mass number ———— 12
Atomic number ———— 6

$$^{12}_{6}C$$

Mass number $= 12$
Atomic number $= 6$
No of protons $= 6$
No of electrons $= 6$
No of neutrons $= 12 - 6 = 6$

Mass number ——— 207
Atomic number ——— 82

$$^{207}_{82}Pb$$

Mass number $=$ ____
Atomic number $=$ ____
No of protons $=$ ____
No of electrons $=$ ____
No of neutrons $=$ ____

Copy and complete the table below

Element	Symbol	No.of electrons	No.of protons	No.of neutrons	Atomic no.	Mass no.
Hydrogen	H	1	**1**	0	**1**	1
Magnesium	Mg	12	12	12	12	24
Silicon	Si	14		14		
Sulphur	S	16				32
Chlorine	Cl				17	35
Argon	Ar	18				40
Potassium	K	19	19	20		
Calcium	Ca			20		40

Isotopy, isotopes and calculation of relative atomic mass (R.A.M.)

Isotopy is the existence of atoms of the same element with the same atomic number but different mass number.

Isotopes are atoms of the same element with the same atomic number but different mass numbers.(Atoms of the same element with the same number of protons but different number of neutrons). Examples of elements with isotopes are chlorine carbon etc.

Examples of isotopic elements

Element	Isotope	Isotope	Atomic no.	Protons	Neutrons	Mass number
Carbon	Carbon – 12	${}^{12}_{6}C$				
	Carbon – 14	${}^{14}_{6}C$				
Silver	Silver – 107	${}^{107}_{47}Ag$				
	Silver – 109	${}^{109}_{47}Ag$				
Chlorine	Chlorine – 35	${}^{35}_{17}Cl$				
	Chlorine – 37	${}^{37}_{17}Cl$				

Relative atomic mass (R.A.M)

R.A.M – This is the average mass of an element compared with an atom of a single carbon – 12, the mass of which is 12 . When calculating the RAM of an element, it can be a fraction because it is an average mass of the isotopes of the element. R.A.M is measured accurately using a **mass spectrometer** and the results obtained from such an instrument are known as **mass spectrograms**. Mass spectrometer can also be used to obtain the relative abudance of isotopes. R.A.M is always close to the most abundant isotopes because the most abundant contributes more to the R.A.M.

Example 1

Chlorine is made up of two isotopes as follows:

1. Cl – 35 (75%)
2. Cl – 37 (25%)

Formula for calculating the R.A.M.

$$\boxed{\textbf{R.A.M} = \frac{P_1M_1 + P_2M_2}{100}} \quad \text{or} \quad \boxed{\textbf{R.A.M} = \frac{R_1M_1 + R_2M_2}{R_1 + R_2}}$$

P_1 = Percentage of isotope one — P_2 = Percentage of isotope two

M_1 = Mass of isotope one — M_2 = Mass of isotope two

R_1 = Ratio of isotope one — R_2 = Ratio of isotope two

$$\text{R.A.M of chlorine} = \frac{P_1M_1 + P_2M_2}{100} = \frac{(35\times 75)+(37\times 25)}{100} = \frac{2625+925}{100} = 35.5$$

Example 2

Chlorine is made up of two isotopes Cl – 37 and Cl – 35 which occur in the ratio 1:3 respectively. Calculate R.A.M of chlorine.

$$\text{R.A.M of chlorine} = \frac{R_1M_1 + R_2M_2}{R_1 + R_2} = \frac{(1\times 37)+(3\times 35)}{1+3} = \frac{37+105}{4} = \frac{142}{4} = 35.5$$

Example 3

The R.A.M of chlorine is 35.5. Calculate the relative abundance of its two isotopes; Cl – 37 and Cl – 35.

Solution:

Let percentage of Cl – 37 be $x\%$. That of Cl – 35 will be $(100-x)\%$. This is because % mass sums up to 100%.

$$\textbf{R.A.M} = \frac{P_1M_1 + P_2M_2}{100}$$

Substituting:

$$35.5 = \frac{(35\times x)+37(100-x)}{100}$$

$$35.5\times 100 = 35x + 3700 - 37x$$

$$\Rightarrow 3550 = 35x - 37x + 3700$$

$$37x - 35x = 3700 - 3550$$

$$2x = 150 \Rightarrow x = 75\%$$

$$\therefore\ \text{Cl} - 37 = 75\%$$

$$\text{Cl} - 35 = (100-75) = 25\%$$

NB: The R.A.M. of isotopic elements is always close to the most abundant isotope.

Question 1

Chlorine has two isotopes naturally occurring; Cl – 35 and Cl – 37. Given that 75% of all naturally occurring chlorine is chlorine – 35:

(a) What is the percentage abundance of the other isotope?

(b) Out of 6000 atoms of chlorine, how many will be of the heavier isotope?

(c) Calculate the R.A.M. of chlorine.

Question 2

An element X has R.A.M of 20.2 and has two isotopes $X-20$ and $X-21$.

(a) Which is the most abundant isotope?

(b) Calculate the % of the most abundant isotope.

Question 3

Boron has two isotopes $B-11$ and $B-10$ which occur in the ratio **4:1**. Calculate the R.A.M of boron.

Question 4

Carbon has two isotopes; $C-14$ and $C-12$, and its relative atomic mass is **12.01**. Calculate the relative abudance of its isotopes.

Question 5

Calculate the R.A.M of potassium from the isotopic compositions given below.

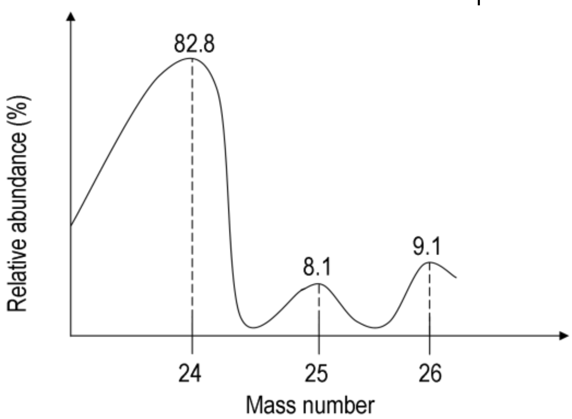
Isotope	**K – 39**	**K – 40**	**K – 41**
relative abudance	93.1	0.01	6.89

Question 6

Element M has two isotopes with mass numbers **25** and **29** respectively. The R.A.M. of E is **28.078**. Determine the percentage abundances of the isotopes.

Question 7

The graph below relates to element Q. Use it to determine the relative atomic mass (R.A.M.) of Q show your working. The height of each peak is proportional to the relative abundance of each isotope.



Question 8

Element **M** has two isotopes with mass numbers **25** and **29** respectively. The R.A.M. of

M is **28.078**. Determine the percentage abundances of the isotopes.

Question 9

A neutral atom of silicon contains **14** electrons, **92%** of silicon – **28**, **5%** silicon – **29** and **3%** of silicon – **30**. Calculate the R.A.M of silicon.

Question 10

Magnesium has three isotopes; **Mg – 24**, **Mg – 25** and **Mg – 26** with relative abundances of **79%**,**10%** and **11%** respectively. Calculate the relative atomic mass of magnesium.

ENERGY LEVEL AND ELECTRON ARRANGEMENT

Energy level – these are the regions around the nucleus that are occupied by electrons. The arrangement of electrons in the energy levels is called electron arrangement. The energy level are numbered 1,2,3, starting with the one closest to the nucleus. Each energy level can only accommodate a given maximum number of electrons

Energy level	Maximum electrons
1st energy level	2
2nd energy level	8
3rd energy level	8

NB: The above will only apply for the first 20 elements.
Electron arrangement is also called electronic configuration.
E.g. for lithium with 3 electrons; 2.1
Boron with 5 electrons; 2.3
It can also be represented by a diagram using either a cross $(\times)$ or a dot $(\bullet)$ to represent an electron.

E.g **Na** 11 = 2.8.1, Silicon 14 = 2.8.4
Write the configuration of the first 20 elements, and use it to determine the group, period, whether the element is a metal or a non – metal and the valence.
Note:
The columns for **group, period, metal or non-metal and valence should only be filled when the student has been taught the subtopic of periodic table and valence respectively**

Mnemonic	Element	Atomic number	Symbol	Electron arrangement	Group	Period	Metal/non metal	Valency
Hi	Hydrogen	1						
Helen	Helium	2						
Listens	Lithium	3	Li	2.1	1	2	Metal	1
B	Beryllium	4						
B	Boron	5						
C	Carbon	6						
News	Nitrogen	7						
On	Oxygen	8						
Friday	Fluorine	9	F	2.7	7	2	Non metal	1
Night	Neon	10						
Somali	Sodium	11						
Men	Magnesium	12						
Arrested	Aluminium	13						
Six	Silicon	14						
People	Phosphorous	15						
Suspected of	Sulphur	16						
Committing	Chlorine	17						
A	Argon	18						
Political	Potassium	19						
Crime	Calcium	20	Ca	2.8.8.2	2	4	Metal	2

PERIODIC TABLE

This is a table of elements are arranged according to their atomic numbers

It is made of periods which run across/rows and groups which run vertically/columns

Periodic table has eight major groups and seven periods

A section of periodic table showing the first twenty elements is shown below

For The First 20 Elements

Groups

Period	I	II		III	IV	V	VI	VII	VIII
1	H								He
2	Li	Be		B	C	N	O	F	Ne
3	Na	Mg		Al	Si	P	S	Cl	Ar
4	K	Ca							

hydrogen 1 **H** 1.0079																		helium 2 **He** 4.0026
lithium 3 **Li** 6.941	beryllium 4 **Be** 9.0122												boron 5 **B** 10.811	carbon 6 **C** 12.011	nitrogen 7 **N** 14.007	oxygen 8 **O** 15.999	fluorine 9 **F** 18.998	neon 10 **Ne** 20.180
sodium 11 **Na** 22.990	magnesium 12 **Mg** 24.305												aluminium 13 **Al** 26.982	silicon 14 **Si** 28.086	phosphorus 15 **P** 30.974	sulfur 16 **S** 32.065	chlorine 17 **Cl** 35.453	argon 18 **Ar** 39.948
potassium 19 **K** 39.098	calcium 20 **Ca** 40.078		scandium 21 **Sc** 44.956	titanium 22 **Ti** 47.867	vanadium 23 **V** 50.942	chromium 24 **Cr** 51.996	manganese 25 **Mn** 54.938	iron 26 **Fe** 55.845	cobalt 27 **Co** 58.933	nickel 28 **Ni** 58.693	copper 29 **Cu** 63.546	zinc 30 **Zn** 65.39	gallium 31 **Ga** 69.723	germanium 32 **Ge** 72.61	arsenic 33 **As** 74.922	selenium 34 **Se** 78.96	bromine 35 **Br** 79.904	krypton 36 **Kr** 83.80
rubidium 37 **Rb** 85.468	strontium 38 **Sr** 87.62		yttrium 39 **Y** 88.906	zirconium 40 **Zr** 91.224	niobium 41 **Nb** 92.906	molybdenum 42 **Mo** 95.94	technetium 43 **Tc** [98]	ruthenium 44 **Ru** 101.07	rhodium 45 **Rh** 102.91	palladium 46 **Pd** 106.42	silver 47 **Ag** 107.87	cadmium 48 **Cd** 112.41	indium 49 **In** 114.82	tin 50 **Sn** 118.71	antimony 51 **Sb** 121.76	tellurium 52 **Te** 127.60	iodine 53 **I** 126.90	xenon 54 **Xe** 131.29
caesium 55 **Cs** 132.91	barium 56 **Ba** 137.33	57-70 *	lutetium 71 **Lu** 174.97	hafnium 72 **Hf** 178.49	tantalum 73 **Ta** 180.95	tungsten 74 **W** 183.84	rhenium 75 **Re** 186.21	osmium 76 **Os** 190.23	iridium 77 **Ir** 192.22	platinum 78 **Pt** 195.08	gold 79 **Au** 196.97	mercury 80 **Hg** 200.59	thallium 81 **Tl** 204.38	lead 82 **Pb** 207.2	bismuth 83 **Bi** 208.98	polonium 84 **Po** [209]	astatine 85 **At** [210]	radon 86 **Rn** [222]
francium 87 **Fr** [223]	radium 88 **Ra** [226]	89-102 **	lawrencium 103 **Lr** [262]	rutherfordium 104 **Rf** [261]	dubnium 105 **Db** [262]	seaborgium 106 **Sg** [266]	bohrium 107 **Bh** [264]	hassium 108 **Hs** [269]	meitnerium 109 **Mt** [268]	ununnilium 110 **Uun** [271]	unununium 111 **Uuu** [272]	ununbium 112 **Uub** [277]		ununquadium 114 **Uuq** [289]				

*Lanthanide series	lanthanum 57 **La** 138.91	cerium 58 **Ce** 140.12	praseodymium 59 **Pr** 140.91	neodymium 60 **Nd** 144.24	promethium 61 **Pm** [145]	samarium 62 **Sm** 150.36	europium 63 **Eu** 151.96	gadolinium 64 **Gd** 157.25	terbium 65 **Tb** 158.93	dysprosium 66 **Dy** 162.50	holmium 67 **Ho** 164.93	erbium 68 **Er** 167.26	thulium 69 **Tm** 168.93	ytterbium 70 **Yb** 173.04
**Actinide series	actinium 89 **Ac** [227]	thorium 90 **Th** 232.04	protactinium 91 **Pa** 231.04	uranium 92 **U** 238.03	neptunium 93 **Np** [237]	plutonium 94 **Pu** [244]	americium 95 **Am** [243]	curium 96 **Cm** [247]	berkelium 97 **Bk** [247]	californium 98 **Cf** [251]	einsteinium 99 **Es** [252]	fermium 100 **Fm** [257]	mendelevium 101 **Md** [258]	nobelium 102 **No** [259]

N/B: Groups are indicated using capital roman numbers in brackets.

- Elements with the same number of electrons in the outermost energy level are in the same **group.**
- Elements with the same number of occupied energy levels are in the same **period**.
- **Elements whose outermost energy level are filled up are in group eight and that is why Helium is in group eight and not group two.**
- Elements in group one to three are usually metals while from group four to eight comprise of non –metals
- Silicon is a **metalloid (semi-metal)** and hydrogen is a non-metal although it is in group one.

- Between group 2 and 3 it is a group of elements called **transition elements** which have unique characteristics e.g.
  - They have variable valencies e.g. copper has valency of 1 or 2.
  - They form coloured compounds.
  - They have high melting points and boiling points.
  - Some form complex ions.
  - They are good conductors of heat and electricity e.g. Cu ,Fe ,Pb ,Zn.

Note
- ✓ Hydrogen is a non-metal though in group (I).
- ✓ Helium is in group 8 and not group 2 as its outermost energy level is filled up.

ION FORMATION

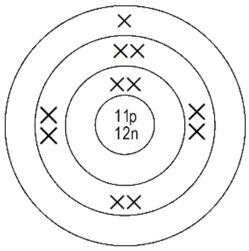
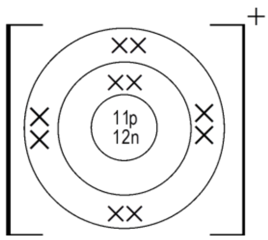
Elements in group eight are called noble gases, this is because they have stable configuration and don't react with other elements and exist as single atoms. Atoms of other elements try to achieve the noble gas configuration by either losing or gaining electrons. Atoms which have lost or gained electrons consequently have a charge.

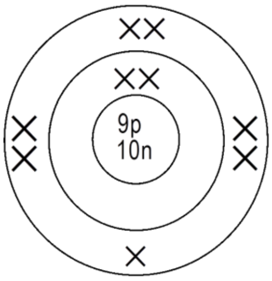
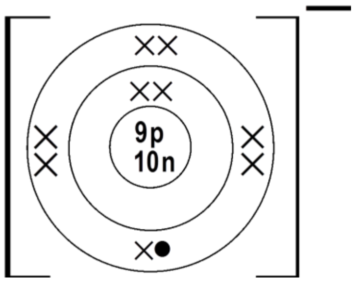
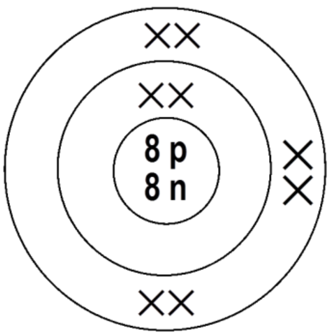
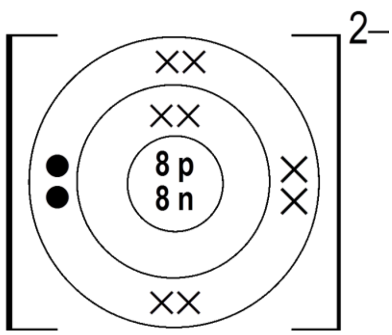
A **charged atom** is called an **ion.** A **positively charged ion** is called a **cation**, while a negatively charged ion is called an **anion**. When element loses electrons, they acquire a positive charge (cation). When an element gains electrons, there is increase in negatively charged particles hence acquire a negative charge (anion).

Atoms with ionic configuration of 2, 2.8, 2.8.8 are said to be stable.

E.g. Magnesium with atomic number of 12 has configuration of 2.8.2. It can become stable by either losing two electrons to become 2.8 or gaining 6 to became 2.8.8. But it usually loses 2 electron as this requires less energy than gaining six electrons.

NB: When drawing the structure of atom or ion, it is important to show the composition of the nucleus. Electrons are shown using dots $(\bullet)$ or crosses $(\times)$.

Element	Atom structure	Ion structure and configuration
Sodium (Na) $^{23}_{11}Na$	**Na—2.8.1 loses one electron to become stable.** 	$Na^{+}-2.8$ 
Beryllium $^{7}_{4}Be$		

Magnesium $^{24}_{12}Mg$		
Fluorine $^{19}_{9}F$	**Fluorine atom 2.7 gains one electron to become stable.** 	**Fluoride ion –** $F^{-} - 2.8$ **the dot (•) represent the gained electron.** 
Phosphorous $^{31}_{15}P$		
Calcium $^{40}_{20}Ca$		
Oxygen $^{16}_{8}O$		

Electronegativity

This is the tendency of an atom to gain electrons. Electronegativity increases across the period but decreases down the group. This is because as you move across the period there is an increase in nuclear charge and thus a greater attraction for any incoming electron, but as you go down a group, there is an increase in the number of energy levels and the shielding effect of the orbitals makes it harder to gain electrons.

Electropositivity

This is the tendency of an atom to lose electrons. It decreases across the period but increases down the group; this is because as you move across the period, there is decrease in atomic radius and increase in nuclear charge and it will be harder to lose electron. **Electropositivity is also called metallic character.**

As you go down the group, there is an increase in the number of occupied energy levels increasing the shielding effect which makes it easier to loose electrons.

VALENCY

Valency is the combining power of an element. It is ***usually equal to the number of electrons lost or gained***.

E.g. Sodium loses one electron to became stable hence has valency 1. Fluorine gains one electron to became stable and hence valency one.

- ✓ Group (I) and group (VII) elements lose and gain one electron respectively and hence have **valency one**.
- ✓ Group (II) and group (VI) loses and gains two electrons and hence has a **valency of two.**
- ✓ Group (III) and group (V) loses and gains three electrons hence has a **valency of three**.
- ✓ Group (IV) has a **valency of four**.
- ✓ Group (VIII) have **valency 0** because they have a stable configuration and therefore don't lose or gain electrons.

Valency 1	Valency 2	Valency 3	Valency 4
H, Cl, F	Mg, Ca, Be, Ba	Al	C
Br, Na, K	Fe, Cu, Hg	B	Si
Li, Ag, Cu, N	Pb, Zn, O, N	N, Fe, N	Pb, N

Oxidation number/ oxidation state

This is the apparent charge of a particle. Atom in neutral state has an oxidation of zero.

Particle	Oxidation number
Na^+	+1
Mg^{2+}	+2

Al^{3+}	+3
Cl^-	–1
S^{2-}	–2
Br	O

Radicals

- ✓ These are groups of atoms that react together as a single unit.
- ✓ They are usually charged and their valency is equivalent to their charge.

Valency 1	Valency 2	Valency 3
Ammonium (NH_4^+)	Sulphate (SO_4^{2-})	Phosphate (PO_4^{3-})
Chlorate (ClO_3^-)	Carbonate (CO_3^{2-})	
Acetate /ethanoate (CH_3COO^-)	Sulphite (SO_3^{2-})	
Nitrate (NO_3^-), Nitrite (NO_2^-)	Hydrogen phosphate (HPO_4^{2-})	
Hydrogen carbonate (HCO_3^-)	Peroxide (O_2^{2-})	
Hypochlorite (OCl^-) ,Hydroxide (OH^-)		
Hydrogen sulphate (HSO_4^-)		
Perchlorate (ClO_4^-)		

WRITING CHEMICAL FORMULAE

When writing chemical formula of a compound you need to know the chemical symbol of the combining elements and their valencies

Steps to follow

1) Write the chemical symbols of the combining elements/radicals
2) Crisscross their valencies
3. Incase the compound has more than one radical the number of radicals is indicated using brackets. Refer to formular of Magensium nitrate example 4

Compound	**Working**	**Chemical formula**
Magnesium Nitride	Mg^2N^3 crossmulply $\Rightarrow$ valencies	Mg_3N_2
Sodium Oxide	Na^1O^2 crossmulply $\Rightarrow$ valencies	Na_2O
Magnesium Sulphate	$Mg^2SO_4^{valence 2}$ divide by $2 \Rightarrow$	$MgSO_4$
Magnesium Nitrate	$Mg^2 \; NO_3^1$	$Mg(NO_3)_2$
Ammonim Chloride		
Magnesium Chloride		
Sodium Phosphate		
Calcium Carbonate		

Calcium Phosphate		
Sodium Sulphate		
Ammonium phosphate	$NH_4^1 \; {}^3PO_4$	$(NH_4)_3PO_4$
Ammonium Sulphate		
Aluminium Chloride		
Aluminium Sulphate		
Aluminium Carbonate		
Aluminium Phosphate		
Aluminium Nitride		

Aluminium Sulphide		
Aluminium Nitrate		
Magnesium Hydroxide		
Magnesium Phosphate		
Magnesium Sulphide		
Silver Chloride		
Sliver Carbonate		
Zinc Nitrate		
Zinc Carbonate		

Zinc Phosphate		
Mercury Nitrate		
Mercury Sulphate		
Calcium Chlorate		
Magnesium Per Chlorate		
Sodium Chlorite		
Potassium Chlorate		
Sodium Sulphite		
Sodium hydrogen carbonate		
Potassium hydrogen phosphate		

For elements with more variab**le** valences e.g Fe(2or3) Pb(2 0r 4) Cu(1 or 2) the valency used is indicated using roman numbers in brackets when naming their compound Eg

COMPOUND	**WORKING**	**FORMULA**
Carbon (II) oxide	$C^{\frac{2}{2}=1}$ $O^{\frac{2}{2}=1}$	CO
Carbon (IV) oxide	$C^{\frac{4}{2}=2}$ $O^{\frac{2}{2}=1}$	CO_2
Sulphur (IV) oxide		
Sulphur (VI) oxide		
Chromium (III) oxide		
Chromium (VI) oxide		
Lead (II) sulphate		
Lead (IV) chloride		
Lead (IV) oxide		
Lead (II) oxide		
Nitrogen (I) oxide		
Nitrogen(II) oxide		
Nitrogen (IV) oxide		

Silicon (IV) chloride		
Iron (II) chloride		
Iron (III) chloride		
Iron (II) sulphate		
Iron (III) sulphate		
Iron (II) nitrate		
Iron (III) nitrate		
Iron (II) oxide		
Iron (III) oxide		
Iron (II) phosphate		
Iron (III) phosphate		
Lead (II) phosphate		
Lead (IV) phosphate		

Lead (II) carbonate		
Lead (IV) carbonate		
Copper (II) oxide		

Copper (I) oxide		
Copper (II) Sulphate		
Lead (II) ethanoate		

WRITING AND BALANCING CHEMICAL EQUATIONS

When writing chemical equation you need to know the

a)Reactants and the products

b) Chemical symbol of reactants and products

Steps to follow

a) Write the word equation

$$Sodium\ Hydroxide\ + sulphuric\ (VI)\, acid \longrightarrow Sodium\ Sulphate\ + water$$

b) Write using chemical symbols

$$NaOH\ +\ H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$$

c) balance the equation

$$2NaOH\ +\ H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$$

d) include the state symbols

$$NaOH_{(aq)} +\ H_2SO_{4(aq)} \longrightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$$

WRITE THE FOLLOWING WORD EQUATIONS INTO CHEMICAL EQUATIONS AND BALANCE THEM, (PUT THE STATES)

1. $Iron + Sulphur \longrightarrow Iron\ (II)\ Sulphide$

2. Hydrogen + Chlorine gas $\longrightarrow$ Hydrogen Chloride

3. Hydrogen + Nitrogen $\longrightarrow$ Ammonia

4. $Sulphur + Oxygen \longrightarrow Sulphur\ (IV)\ Oxide$

5. $Carbon + Oxygen \longrightarrow Carbon\ (IV) Oxide$

6. $Zinc + Sulphuric (VI)\ acid \longrightarrow Zinc\ Sulphate + hydrogen$

7. $Magnesium + Hydrochloric\ acid \longrightarrow Magnesium\ Chloride + Hydrogen$

8. $Aluminium + Hydrochloric\ acid \longrightarrow Aluminium\ chloride + Hydrogen$

9. $Sodium\ Carbonate + Hydrochloric\ acid \longrightarrow Sodium\ Chloride + Water + Carbon\ (IV)\ Oxide$

10. $Calcium\ Carbonate + Hydrochloric\ acid \longrightarrow Calcium\ Chloride + Water + Carbon\ (IV)\ Oxide$

11. $Copper(II)Carbonate + Hydrochloric\ acid \longrightarrow Copper(II)Chloride + Water + Carbon\ (IV)\ Oxide$

12. $Sodium$ Hydrogen $Carbonate + Hydrochloric\ acid \longrightarrow Sodium\ Chloride + Water + Carbon\ (IV)\ Oxide$

13. $Calcium\ Carbonate + Sulphuric\ (VI) acid \longrightarrow Calcium Sulphate + Water + Carbon\ (IV)\ Oxide$

14. $Sodium\ hydroxide + Hydrochloric\, acid \longrightarrow Aluminium chloride + Water$

15. $Sodium\ hydroxide + Carbon\left(IV\right)Oxide \longrightarrow sodium\ carbonate + Water$

16. $Sodium\ peroxide + Water \longrightarrow sodium\ hydroxide + oxygen\ gas$

17. $Alu\min iumOxide + Nitric\ \left(IV\right)\ acid \longrightarrow Alu\min iumnitrate + water$

18. $Zinc\, Oxide + Hydrochloric\, acid \longrightarrow Zinc\, Chloride + Water$

19. $Copper\left(II\right)\ Oxide + Sulphuric\ \left(IV\right)\ acid \longrightarrow Copper\ \left(II\right)\ Sulphate + Water$

20. $Copper + Oxygen \longrightarrow Copper\ \left(II\right)\ Oxide$

21. $Phosphorus + Oxygen \longrightarrow Phosphorus\left(III\right)Oxide$

22. $Phosphorus + Oxygen \longrightarrow Phosphorus\left(V\right)Oxide$

23. $Iron + Oxygen \longrightarrow Iron\left(III\right)Oxide$

24. $Magnesium + Oxygen \longrightarrow Magnesium\, Oxide$

25. $Magnesium + Nitrogen \longrightarrow Magnesium\ nitride$

25. $Carbon + Oxygen \longrightarrow Carbon\left(II\right)Oxide$

26. $Carbon + Copper\ (II)\ Oxide \longrightarrow Copper + Carbon\ (IV) Oxide$

27. $Sodium + Water \longrightarrow Sodium\ hydroxide + Hydrogen$

28. $Calcium + Water \longrightarrow Calcium\ hydroxide + Hydrogen$

29. $Potassium + Water \longrightarrow Potassium\ hydroxide + Hydrogen$

30. $Copper(II)Oxide + Hydrogen \longrightarrow Copper + Water$

31. $Lead(II)Oxide + Hydrogen \longrightarrow Lead + Water$

32. $Iron(III)Oxide + carbon \longrightarrow Iron + Carbon\ (IV)\ oxide$

33. $Hydrogen + Oxygen \longrightarrow Water$

34. $Lead(II)\ nitrate + Sodium\ Chloride \longrightarrow Lead\ (II)\ Chloride + Sodium\ nitrate$

35. $Chlorine + water \longrightarrow Hydrochloric\ acid + Chloric(I)\ acid$

36. $Magnesium + Water \longrightarrow Magnesium\ Hydroxide + Hydrogen$

36. $Lead(II)nitrate + Sodium\ Chloride \longrightarrow Lead\ (II)Chloride + Sodium\ nitrate$

37. $Barium\ Chloride + Sodium\ Carbonate \longrightarrow Barium\ Carbonate + Sodium\ Chloride$

38. $Lead\ (II)\ nitrate + Potassium\,iodide \longrightarrow Lead\ (II)\ Iodide + Potassium\,nitrate$

39. $Iron + Copper(II)\ Sulphate \longrightarrow Iron\ (II)\ Sulphate + Copper$

40. $Sodium\ hydroxide + chlorine\,gas \longrightarrow Sodium\,chlorate(V) + sodium\,chloride + Water$

41. $Aluminium\ Oxide + Hydrochloric\ acid \longrightarrow Aluminium\ Chloride + Water$

42. $Aluminium\ Oxide + Sodium\ hydroxide \longrightarrow Sodium\ aluminate + Water$

43. $Silver\ Chloride + Water \longrightarrow Silver\ Oxide + Hydrochloric\ acid$

44. $Phosphorus\ (V)chloride + Water \longrightarrow Phosphoric\ (V)acid + Hydrochloric\ acid$

45. $Iron + Chlorine \longrightarrow Iron\ (III)\ Chloride$

46. $Iron + Hydrogen\ Chloride \longrightarrow Iron\ (II)Chloride + hydrogen\ gas$

47. $Lithium + Chlorine \longrightarrow Lithium\ Chloride$

48. $Lithium + Water \longrightarrow Lithium\ hydroxide + Hydrogen$

49. $Lead(II)Oxide + Carbon \longrightarrow Lead + Carbon(IV)\ Oxide$

50. $Iron(III)\ Oxide + Carbon \longrightarrow Iron + Carbon(IV)Oxide$

51. $Calcium\ hydroxide + Carbon\ (IV)\ Oxide \longrightarrow Calcium\ Carbonate + Water$

52. $Carbon\ (IV)\ Oxide\ +\ Water \longrightarrow Carbonic\ acid$

53. $Calcium\ Carbonate + Carbon\ (IV)\ Oxide + Water \longrightarrow Calcium\ hydrogen\ Carbonate$

54. $Magnesium + Carbon\ (IV)\ Oxide \longrightarrow Magnesium\ Oxide + Carbon$

55. $Iron\ (III)\ Oxide\ +\ Carbon\ (II)\ Oxide \longrightarrow Iron\ +\ Carbon\ (IV)\ Oxide$

56. $\underset{carbonate}{Calcium\ hydrogen} + \underset{acid}{Hydrochloric} \longrightarrow \underset{chloride}{Calcium} + \underset{oxide}{Carbon\ (IV)} + Water$

57. $Sodium\ hydrogen\ Carbonate \xrightarrow{Heat} Sodium\ Carbonate + Water + Carbon\ (IV)\ Oxide$

58. $hydrogen \xrightarrow{manganese\ (IV)\ oxide} water\ +\ oxygen\ gas$

59. $Lead\ Carbonate \xrightarrow{Heat} Lead\ (II)\ Oxide\ +\ Carbon(IV)\ Oxide$

60 $Ammonium\ Carbonate \xrightarrow{Heat} Ammonia + Water + Carbon\ (IV)\ Oxide$

61. $\underset{carbonate}{Ammonium\ hydrogen} + Sodium\ Chloride \longrightarrow \underset{carbonate}{Sodium\ hydrogen} + \underset{chloride}{Ammonium}$

62. $Calcium\ Hydroxide\ +\ Ammonium\ Chloride \longrightarrow Calcium\ Chloride\ +\ Ammonium + Water$

63. $Potassium\ hydroxide + Carbon\ (IV)\ Oxide \longrightarrow Potassium\ hydrogen\ Carbonate$

64. Ammonium nitrite $\xrightarrow{Heat}$ Water + Nitrogen gas

65. Calcium hydroxide + Ammonium chloride $\longrightarrow$ Calcium chloride + Water + Ammonia gas

66. Hydrochloric acid + Ammonia gas $\longrightarrow$ Ammonium Chloride

67. Sulphuric (VI) acid + Ammonia gas $\longrightarrow$ Ammonium Sulphate

68. Sulphuric (IV) acid + Ammonium hydroxide $\longrightarrow$ Ammonium Sulphate + Water

69. Hydrochloric acid + Ammonium hydroxide $\longrightarrow$ Ammonium Chloride + Water

70. Nitric (V) acid + Ammonium hydroxide $\longrightarrow$ Ammonium nitrate + Water

71. Ammonia gas + Oxygen $\longrightarrow$ Water + Nitrogen

72. Nitrogen (IV) Oxide + Oxygen + Water $\longrightarrow$ Nitric (V) acid

73. Copper (II) Oxide + Ammonia gas $\longrightarrow$ Copper + Water + Nitrogen

74. Copper + concentrated Nitric (V) acid $\longrightarrow$ Copper (II) Nitrate + Water + Nitrogen (IV) oxide

75. Nitrogen (II) Oxide + Oxygen $\longrightarrow$ Nitrogen (IV) Oxide

76. Lead (II) nitrate $\xrightarrow{Heat}$ Lead (II) Oxide + Nitrogen (IV) Oxide + Oxygen

77. $Magnesium \quad + \quad Nitrogen\ (IV)\ Oxide \longrightarrow Magnesium\ Oxide + \ Nitrogen$

78. $Sulphuric\ (VI)\ Acid + \quad Potassium\ nitrate \longrightarrow \ Potassium\ hydrogen\ Sulphate \ + \ Nitric\ (V)\ acid$

80. $Copper + \ Nitric\ (V)\ acid \longrightarrow Copper\ nitrate + Water \ + Nitrogen\ (II)\ Oxide$

81. $Sodium\ nitrate \xrightarrow{Heat} \ Sodium\ nitrite \quad + \quad Oxygen$

82. $Potassium\ nitrate \xrightarrow{Heat} Potassium\ nitrite + \quad Oxygen$

83. $Copper\ (II)\ nitrate \xrightarrow{Heat} Copper\ (II)\ Oxide + Nitrogen\ (IV)\ Oxide + Oxygen$

84. $Silver\ nitrate \xrightarrow{Heat} Silver \ + \ Nitrogen\ (IV)\ Oxide \ + \ Oxygen$

85. $Sulphur + \ Nitric\ (V)\ acid \longrightarrow Sulphuric\ (VI)\ acid + Nitrogen\ (IV)\ Oxide + Water$

86. $Sodium\,Sulphite \ + \quad Hydrochloric\,acid \longrightarrow Sulphur\ (IV) Oxide \ + \quad Sodium\,Chloride \quad + \ Water$

87. $Copper \quad + \ Sulphuricacid \quad \longrightarrow Copper\ (II)\ Sulphate \ + Water \ + Sulphur\ (IV) \ Oxide$

88. $Nitric\ (V)\ acid \ + \ Sulphur\ (IV)\ Oxide \longrightarrow Nitrogen\ (IV)\ Oxide + Sulphuric\ (IV)\ acid$

89. $Magnesium \ + \quad Sulphur\ (IV)\ Oxide \longrightarrow Magnesium\ Oxide + Sulphur$

90. $Calcium\ hydroxide \ + Sulphur\ (IV)\ Oxide \longrightarrow Calcium\ Sulphate \ + \ Water$

91. $Hydrogen\ sulphide\ +\ Sulphur\ (IV)\ Oxide \longrightarrow\ Sulphur\ +water$

92. $Ammonia\ gas\ +\ oxygen \longrightarrow Nitrogen\ (II)\ oxide\ +\ water$

93. $Chlorine\ gas\ +potassium\ iodide \longrightarrow potassium\ chloride\ +\ iodine$

94. $Ammonium\ nitrate \longrightarrow nitrogen\ (I)\ oxide\ +\ water$

95. $Sodium\ Hydroxide + Nitrogen\ (IV) Oxide \longrightarrow\ sodium\ Nitrate\ +\ sodium\ Nitrite\ +Water$

96. $Ammonia + Carbon(IV)oxide \longrightarrow Urea\ +\ water$

97. $Iron + Steam \longrightarrow Tri-Irontetraoxide\ +\ hydrogen\ gas$

98. $Potassium\ Manganate\ (VII) \xrightarrow{heat} Potassium\ Manganate\ (VI) + Manganese\ (IV)\ oxide\ +Oxygen\ gas$

99. $Phosphorous + Nitrogen\ (I) Oxide \longrightarrow Phosphorous\ (V) Oxide + Nitrogen\ gas$

100. $Ammonia + lead(II)oxide \longrightarrow\ lead\ metal + Nitrogen\ gas + water$

Balance the equations below

1. $Na + O_2 \rightarrow Na_2O$

2. $Na + Cl_2 \rightarrow NaCl$

3. $CO + O_2 \rightarrow CO_2$

4. $Mg + O_2 \rightarrow MgO$

5. $CH_4 + O_2 \rightarrow CO_2 + H_2O$

6. $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

7. $CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$

8. $H_2 + Cl_2 \rightarrow HCl$

9. $Na + H_2O \rightarrow NaOH + H_2$

10. $H_2O_2 \rightarrow H_2O + O_2$

11. $AgNO_3 \rightarrow Ag + NO_2 + O_2$

12. $Fe(OH)_3 \rightarrow Fe_2O_3 + H_2O$

13. $Mg + N_2 \rightarrow Mg_3N_2$

14. $Mg_3N_2 + H_2O \rightarrow Mg(OH)_2 + NH_3$

15. $KMnO_4 + HCl \rightarrow KCl + MnCl_2 + Cl_2 + H_2O$

16. $PbO_2 + HCl \rightarrow PbCl_2 + H_2O + Cl_2$

17. $CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$

18. $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$

19. $H_2S + SO_2 \rightarrow S + H_2O$

20. $NO_2 + H_2O \rightarrow HNO_2 + HNO_3$

21. $NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

22. $H_2S + O_2 \rightarrow H_2O + SO_2$

23. $NH_3 + O_2 \rightarrow NO + H_2O$

24. $NH_4NO_3 \rightarrow N_2O + H_2O$

25. $NH_4NO_2 \rightarrow N_2 + H_2O$

26. $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$

27. $PCl_3 + H_2O \rightarrow H_3PO_3 + HCl$

28. $P_2O_5 + H_2O \rightarrow H_3PO_4$

29. $P_2O_3 + H_2O \rightarrow H_3PO_3$

30. $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

31. $H_2CO_3 \rightarrow H_2O + CO_2$

32. $CaCO_3 \rightarrow CaO + CO_2$

33. $CaO + H_2O \rightarrow Ca(OH)_2$

34. $H_2SO_3 \rightarrow H_2O + SO_2$

35. $H_3PO_4 + Ca(OH)_2 \rightarrow CaHPO_4.2H_2O$

36. $SO_3 + H_2O \rightarrow H_2SO_4$

37. $Be(OH)_2 \rightarrow BeO + H_2O$

38. $BaO + H_2O \rightarrow Ba(OH)_2$

39. $Li_2O + H_2O \rightarrow LiOH$

40. $Na_2HPO_4 \rightarrow Na_4P_2O_7 + H_2O$

41. $CaC_2 + N_2 \longrightarrow CaCN_2 + C_2.$

42. $Mg(OH)_2 \longrightarrow MgO + H_2O.$

43. $NaCl + NH_4HCO_3 \longrightarrow NaHCO_3 + NH_4Cl$

44. $Ca(HCO_3)_2 \longrightarrow CaCO_3 + CO_2 + H_2O$

45. $FeS + H_2SO_4 \longrightarrow H_2S + FeSO_4$

46. $(NH_4)_2SO_4 + CaCO_3 \longrightarrow (NH_4)_2CO_3 + CaSO_4$

47. $Hg_2CO_3 \longrightarrow Hg + HgO + CO_2$

48. $BeF_2 + Mg \longrightarrow MgF_2 + Be$

49. $SiO_2 + Ca(OH)_2 \longrightarrow CaSiO_3 + H_2O$

50. $K_2O + H_2O \longrightarrow KOH$

51. $C + H_2O \longrightarrow CO + H_2$

52. $Ca(OH)_2 + CO_2 \longrightarrow Ca(HCO_3)_2$

53. $SiO_2 + Na_2CO_3 \longrightarrow Na_2SiO_3 + CO_2$

54. $BaO_2 + H_2SO_4 \longrightarrow BaSO_4 + H_2O_2$

55. $Na_2Cr_2O_7 + S \longrightarrow Cr_2O_3 + Na_2SO_4$

56. $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$

57. $Fe_2O_3 + SiO_2 \longrightarrow Fe_2Si_2O_7$

58. $CO_2 + NH_3 + H_2O \longrightarrow NH_4HCO_3$

59. $Na_2O + H_2O \longrightarrow NaOH$

60. $NH_4NO_3 \longrightarrow N_2O + H_2O$

61. $Al(OH)_3 + NaOH \longrightarrow NaAlO_2 + H_2O$

62. $H_2O_2 \longrightarrow H_2O + O_2$

63. $Ca(ClO_3)_2 \longrightarrow CaCl_2 + O_2$

64. $PCl_5 + H_2O \longrightarrow POCl_3 + HCl$

65. $Al_2O_3 + Na_2CO_3 \longrightarrow NaAlO_2 + CO_2$

66. $Zn + HCl \longrightarrow ZnCl_2 + H_2$

67. $BeO + C + Cl_2 \longrightarrow BeCl_2 + CO$

68. $BeSO_4 + NH_4OH \longrightarrow Be(OH)_2 + (NH_4)_2SO_4$

69. $K + Br_2 \longrightarrow KBr$

70. $NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$

71. $MnS + HCl \longrightarrow H_2S + MnCl_2$

72. $CaC_2 + H_2O \longrightarrow C_2H_2 + Ca(OH)_2$

73. $CuSO_4 + KCN \longrightarrow Cu(CN)_2 + K_2SO_4$

74. $Ca(OH)_2 + H_3PO_4 \longrightarrow CaHPO_4 + H_2O$

75. $BaCO_3 + HNO_3 \longrightarrow Ba(NO_3)_2 + CO_2 + H_2O_2$

76. $CaO + C \longrightarrow CaC_2 + CO$

77. $Zn(OH)_2 + NaOH \longrightarrow Na_2ZnO_2 + H_2O$

78. $Mn_2O_3 + Al \longrightarrow Al_2O_3 + Mn$

79. $AlN + H_2O \longrightarrow NH_3 + Al(OH)_3$

80.$Ca_3(PO_4)_2 + H_2SO_4 \longrightarrow CaSO_4 + Ca(H_2PO_4)_2$

81. $S + N_2O \longrightarrow SO_2 + N_2$

82. $N_2 + H_2 \longrightarrow NH_3$

83. $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2O + CO_2$

84. $NaOH + Zn(NO_3)_2 \longrightarrow NaNO_3 + Zn(OH)_2$

85. $H_3PO_4 + Ca(OH)_2 \longrightarrow Ca(H_2PO_4)_2 + H_2O$

86.$CaS + H_2O \longrightarrow Ca(HS)_2 + Ca(OH)_2$

87. $Cu + CO_2 + O_2 + H_2O \longrightarrow CuCO_3.Cu(OH)_2$

88. $Sn(OH)_2 + NaOH \longrightarrow Na_2SnO_2 + H_2O$

89. $NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + HCl$

90. $Fe(OH)_3 \longrightarrow Fe_2O_3 + H_2O$

91. $NaOH + Cl_2 \longrightarrow NaCl + NaClO + H_2O$

92. $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

93. $SiH_4 + O_2 \longrightarrow SiO_2 + H_2O$

94. $Pb(OH)_2 + NaOH \longrightarrow Na_2PbO_2 + H_2O$

95. $Si + NaOH + H_2O \longrightarrow Na_2SiO_3 + H_2$

96. $Si + S_8 \longrightarrow Si_2S_4$

97. $CaS_2 + O_2 \longrightarrow CaS_2O_3$

98. $Na_2S_2 + O_2 \longrightarrow Na_2S_2O_3$

99. $HCl + K_2CO_3 \longrightarrow KCl + H_2O + CO_2$

100. $KClO_3 \longrightarrow KCl + O_2$

101. $Zn + NaOH + H_2O \longrightarrow Na_2Zn(OH)_4 + H_2$

102. $Na_2CO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$

103. $Ca(OH)_2 + P_4O_{10} + H_2O \longrightarrow Ca(H_2PO_4)_2$

104. $CaS + H_2O + CO_2 \longrightarrow Ca(HCO_3)_2 + H_2S$

105. $Na + H_2O \longrightarrow NaOH + H_2$

109. $Fe_2O_3 + CO \longrightarrow Fe + CO_2$

110. $Pb(NO_3)_2 \longrightarrow PbO + NO_2 + O_2$

111. $Al_2(SO_4)_3 + Ca(OH)_2 \longrightarrow CaSO_4 + Al(OH)_3$

112 . $Ca_3(PO_4)_2 + H_2SO_4 \longrightarrow CaSO_4 + H_3PO_4$

113. $SiCl_4 + H_2O \longrightarrow H_4SiO_4 + HCl$

114. $Ca + AlCl_3 \longrightarrow CaCl_2 + Al$

115. $FeCl_3 + Ca(OH)_2 \longrightarrow CaCl_2 + Fe(OH)_3$

116. $Al_2O_3 + C + N_2 \longrightarrow AlN + CO$

117. $NO + NaOH \longrightarrow NaNO_2 + H_2O + N_2O$

118. $Pb_3O_4 + HNO_3 \longrightarrow Pb(NO_3)_2 + PbO_2 + H_2O$

119. $KO_2 + CO_2 \longrightarrow K_2CO_3 + O_2$

120. $P_4O_{10} + H_2O \longrightarrow H_3PO_4$

121. $Fe + H_2O + O_2 \longrightarrow Fe_2O_3.H_2O$

122. $H_3PO_4 + HCl \longrightarrow PCl_5 + H_2O$

123. $MnO_2 + KOH + O_2 \longrightarrow K_2MnO_4 + H_2O$

124 $PCl_5 + H_2O \longrightarrow H_3PO_4 + HCl$

125. $P_4O_6 + H_2O \longrightarrow H_3PO_3$

126. $Al\left(OH\right)_3 + H_2SO_4 \longrightarrow Al_2(SO_4)_3 + H_2O$

127. $Fe_2(SO_4)_3 + KOH \longrightarrow K_2SO_4 + Fe\left(OH\right)_3$

128. $Bi(NO_3)_3 + H_2S \longrightarrow Bi_2S_3 + HNO_3$

129. $V_2O_5 + HCl \longrightarrow VOCl_3 + H_2O$

130. $Cr\left(OH\right)_3 + H_2SO_4 \longrightarrow Cr_2(SO_4)_3 + H_2O$

140. $Hg\left(OH\right)_2 + H_3PO_4 \longrightarrow Hg_3(PO_4)_2 + H_2O$

141. $Fe + H_2O \longrightarrow Fe_3O_4 + H_2$

142. $Ca_3P_2 + H_2O \longrightarrow Ca\left(OH\right)_2 + PH_3$

143. $H_2SO_4 + Al\left(OH\right)_3 \longrightarrow Al_2(SO_4)_3 + H_2O$

144. $Al(NO_3)_3 + Na_2CO_3 \longrightarrow Al_2(CO_3)_3 + NaNO_3$

145. $K_2MnO_4 + H_2SO_4 \longrightarrow KMnO_4 + MnO_2 + K_2SO_4 + H_2O$

146. $C_2H_2 + O_2 \longrightarrow CO_2 + H_2O$

147. $C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$

148. $As + NaOH \longrightarrow Na_3AsO_3 + H_2$

149. $H_3BO_3 + Na_2CO_3 \longrightarrow Na_2B_4O_7 + CO_2 + H_2O$

150. $Al + HCl \longrightarrow AlCl_3 + H_2$

151. $V_2O_5 + Ca \longrightarrow CaO + V$

152. $Na_2B_4O_7 + HCl + H_2O \longrightarrow NaCl + H_3BO_3$

153. $C_2H_3Cl + O_2 \longrightarrow CO_2 + H_2O + HCl$

154. $I_2 + HNO_3 \longrightarrow HIO_3 + NO_2 + H_2$

155. $FeS + O_2 \longrightarrow Fe_2O_3 + SO_2$

156 $Ca_3(PO_4)_2 + C \longrightarrow Ca_3P_2 + CO$

157. $NH_3 + O_2 \longrightarrow NO + H_2O$

158. $Hg_2CrO_4 \longrightarrow Cr_2O_3 + Hg + O_2$

159. $S_8 + O_2 \longrightarrow SO_3$

160. $NH_3 + NO \longrightarrow N_2 + H_2O$

161. $HClO_4 + P_4O_{10} \longrightarrow H_3PO_4 + Cl_2O_7$

162. $CO_2 + H_2O \longrightarrow C_6H_{12}O_6 + O_2$

163. $FeS_2 + O_2 \longrightarrow Fe_2O_3 + SO_2$

164. $Si_2H_3 + O_2 \longrightarrow SiO_2 + H_2O$

165 $P_4 + H_2O \longrightarrow H_3PO_4 + H_2$

166. $C_6H_6 + O_2 \longrightarrow CO_2 + H_2O$

167. $C_{10}H_{16} + Cl_2 \longrightarrow C + HCl$

168. $C_7H_6O_2 + O_2 \longrightarrow CO_2 + H_2O$

169. $C_7H_{16} + O_2 \longrightarrow CO_2 + H_2O$

170. $C_7H_{10}N + O_2 \longrightarrow CO_2 + H_2O + NO_2$

171. $KNO_3 + C_{12}H_{22}O_{11} \longrightarrow N_2 + CO_2 + H_2O + K_2CO_3$

REVISION QUESTIONS ON ATOMIC STRUCTURE PERIODIC TABLE

1. An **atom** of sodium is represented as: $^{23}_{11}Na$

(a) Calculate the number of **neutrons** in the nucleus of a sodium atom. [1 mark]

(b) Write the **electronic configuration** of a sodium atom. [1 mark]

(c) State ;

(i) the **group** to which sodium belongs [1 mark]

(ii) the **period** to which sodium belongs [1 mark]

2. Study the grid below which represents part of the **periodic table** and then use it to answer the questions that follow. The letters are not actual symbols of the elements.

A									
	B			C	D			E	F
G								H	

(a) Select two **non-metals** [2 mark]

(b) Select two **metal**s [1 mark]

(c) Select a metalloid

(d) Write the formular of compound formed when the following elements react.

I. G and E

II. B and H

III. C and oxygen

IV. C and E

(e) An element that can form an **ion** with a charge of **+2…………….** [1 mark]

(f) An element that can form an **ion** with a charge of **-1** …………….. [1 mark]

(g) Element P is in period two. and forms ions with the formular P^{3-} place it in the grid above

3. . An element **P** has mass **39** and **20 neutrons**.

a) Calculate the number of **electrons** of the **atom P** [1 mark]

b) **Draw** the **atomic structure** of element **P** using crosses (**x**) to represent electrons in the space provided below [2 marks]

c) Which period does element **P** belong? ………………… [1 mark]

d) An element G consists of isotopes of mass 10 and 11 with some percentage abundances, the relative atomic mass is 10.813, calculate their percentage abundances. (3 marks)

4. (a) What are isotopes? [2 marks]

(c) An element **X** has **3 isotopes** with mass number **22, 24** and **25** in the ratio **2:3:1**. Calculate its **relative atomic mass**. [3 marks]

5. Complete the table below (2½ mks)

Particles	Symbols	Charges	Mass
Electrons			$^{1}/_{1840}$
Neutrons	N		
Proton			

6. An atom of element X can be represented as follows

a) Write the electronic configuration of X. (1 mk)

$$^{24}_{12}X$$

b) Find the number of neutrons in the atom of X. (1 mk)

7. Write the configuration of the ions below (4mks)

a) Ca^{3+}

b) S^{+6}

c) Cl^{+2}

d) S^{2-}

8. A certain atom x has a mass number of 35 and atomic number of 17. What is the number of;

a) Protons (1 mark)

b) Electrons (1 mark)

c) Neutrons (1 mark)

9. An ion P^{2-} has electron arrangement 2.8.
   a) What is the atomic number of the element? (1mk)

   b) Which group and period does the element belong? Give reason. (2mks)

10. A certain element Z has electron arrangement as 2.8.3
    a) What is the formula of the ion. (1 mark)

    b) What is the oxidation number of its ion? (1 mark)

11. Study the table below and answer the question that follows. The letters X, Y and Z are not actual symbols of the element.

Element	Protons	Neutrons	Mass Number
X	12	-	24
Y	17	18	-
Z	-	12	23

   a) Complete the table (3 marks)
   b) Write down the electron arrangement of Z (1 mark)

   c) Write down the formula of the compound formed when X combines with Y. (1 mark)

12. The table below shows some elements in the periodic table. Use it to answer the questions that follow. The letters are not the actual symbols of the elements.

P				T		M	V	U
S		Q						

   a) An element K has atomic number 18. Indicate its position in the grid. (1 mark)
   b) Identify two elements in the same group (2 marks)

c) Identify two elements in the same period. (2 marks)

d) Write the electron configuration of ions of elements: (4 marks)

V Q

P M

13. Write a balanced chemical equation for each of the following compounds.
    a) Action of dilute hydrochloric acid on calcium carbonate. (1mk)

    b) Reacting sodium metal with water to obtain sodium hydroxide and hydrogen gas. (1 mk)

    c) Reaction between Sodium hydroxide and dilute hydrochloric acid. (1 mk)

14. An isotope P has 18 neutrons a mass number of 34.
    a) i). Draw the atomic structure of P. (2 marks)

       ii). Write its electron arrangement. (1 mark)

    b) To which period and group does P belong? Explain your answer. (2 mks)

15. Determine the relative atomic masses of the following elements whose isotopic composition in proportions given. (3mks)

i. Potassium

$^{40}_{19}K$ (0.01%), $^{39}_{19}K$ (93.1%) and $^{41}_{19}K$ (6.89%)

ii. Calculate the number of neutrons in each isotope of potassium. (3 mks)

16. Complete the table below. The symbols are not the actual symbols of elements. (7½ mks)

Atom	Electron arrangement	Formular of Ion	Valency	Oxidation number
X	2.3			
Y	2.8.2			
W	2.7			
Z	2.5			
P	2.8.8			

17. Write the chemical formula of the following. (4 mks)
   a) Sodium chloride

   b) Calcium sulphate

   c) Potassium carbonate

   d) Calcium nitrate

18. The two isotopes of carbon are $^{12}{}_{6}C$ and $^{13}{}_{6}\mathbf{C}$ with relative abundance of 98.8% and 1.2% respectively. Work out the relative atomic mass. (R.A.M) (3 marks)

19. a). What are isotopes? (1 mark)

   b). Lithium has two isotopes $^{7}{}_{3}Li$ and $^{6}{}_{3}Li$ Determine the number of neutrons in $^{7}{}_{3}Li$ (2 marks)

   c). If the relative atomic mass of lithium is 6.94 which of the two isotopes is the most abundant? Give a reason (2 marks)

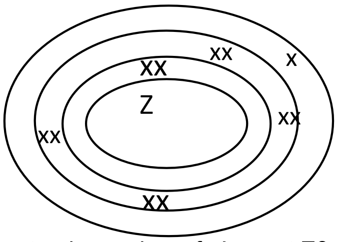
20. Atoms of element X exist as $^{14}{}_{6}X$ and $^{12}{}_{6}X$
   a) What name is given to the two types of atoms? (1 mark)

   b) Use dot (.) and (x) diagrams to illustrates the atomic structure of $^{14}{}_{6}X$ (2 marks)

   c) Write the electron configuration of the atom in (b) hence. Write the formula of the compound formed when it combines with oxygen (0 = 8) (2 mks)

   d)

21. A student represented an atom of element Z as shown in the diagram below



a) What is the atomic number of element Z? Explain (2 marks)

b) Write the formula and electron arrangement of the ion of Z

c) Write equation for the reaction of Z with chlorine

22. An element Y has an electron arrangement of 2.8.5.

a) State the period and group which the elements belongs (1 mark)

b) Write the formula of the most stable ion formed when element Y ionizes (1 mark)

23. An element X, consists of three isotopes with mass number of 22, 24 and 25 with percentage abundance of 89.6%, 6.4% and 4.0% respectively. Find the relative atomic mass of element X, (3 mks)

24. (a) Distinguish between the following:

i. Atomic number and mass number (2 marks)

ii. Mass number and relative atomic mass (2 marks)

(b) Atoms are said to be electrically neutral. Explain (2 mks)

25. A certain atom has the following symbol $^{19}_{9}Y$. What is the number of (5 mks)
    i. Protons
    ii. Neutrons
    iii. Electrons
    iv. Mass number
    v. Atomic number

26. Write the electronic configuration of potassium and carbon with atomic numbers 19 and 6 respectively (2 marks)

27. Define the following:-
    a) Element (1 mark)
    b) Atom (1 mark)
    c) Molecule (1 mark)
    d) Compound (1 mark)

28. An element X is represented as $^{40}_{18}X$. (Note that X does not represent the actual symbol of the element)
    a) Give the number of protons, neutrons and electrons in an atom of the element (1 mark)
    b) Give the electronic configuration of its atom (1 mark)
    c) State the group to which the element belongs (1 mark)
    d) Give a reason for your answer (1 mark)
    e) What would you say about the reactivity of element X? (1 mark)

29. Naturally occurring boron exists as two isotopes $^{10}_{5}B$ with relative abundance of 20% and $^{11}_{5}B$ with a relative abundance of 80%
    a) How many electrons does atom of boron contain (1 mark)
    b) How many neutrons does each atom of the most abundant isotope contain? (1 mark)
    c) Calculate the relative atomic mass of boron (3 marks)

d) Show the distribution of electrons in the energy levels in an atom of $^{11}_{5}B$. Show the composition of the nucleus (2 marks)

30. The following table gives some information about five different atoms. Study it and answer the question that follow. (The letters do not represent the actual symbols at the elements).

ATOM	ATOMIC NUMBER	MASS NUMBER
V	8	16
W	11	23
X	12	26
Y	12	27
Z	17	35

a) Write down the electronic arrangement of

V^{-} (1 mark)
W^{2+} (1 mark)
X^{-2} (1 mark)
Y (1 mark)
Z (1mark)

b) Which of the atoms belong to non-metallic elements? (2 marks)

c) What is the formula for aluminium ion? (1 mark)

31. Write word equations for the reaction between dilute hydrochloric acid and;

a) Magnesium metal (1mk

b) Calcium hydrogen carbonate (1 mark)

c) Potassium hydroxide (1 mark)

32. Write the chemical formulae of the following compounds.

a) Sodium carbonate (1 mark)

b) Zinc chloride (1 mark)

c) Potassium oxide (1 mark)

d) Aluminium sulphate (1 mark)

e) Calcium nitrate (1 mark)

33. An element Z is in group (III) and period 2 in the periodic table
    a) Write down the electron pattern of the atom of element Z. (1 mark)

    b) Draw the atomic structure of the atom of element Z showing the composition of the nucleus. (2 marks)

34. (a) Fill in the blank spaces in the table below.

Particle	Mass Number	Number of Protons	Number of Neutrons	Number of electrons
L1 $^{+}$ ion	-	3	4	-
^{12}C	-	6	-	-

(2 ½ marks)

(b) Write the electron pattern of L1 atom. (1 mark)

(c) Draw the atomic structure of the L1^{+} ion. (1 ½ marks)

35. Write down balanced chemical equations from the reactions below.
    (a) Hydrogen gas reacts with chlorine gas to form hydrogen chloride gas. (2 marks)

    (b) Magnesium carbonate decomposes on heating to form magnesium oxide and carbon (1V) oxide gas. (2 marks)

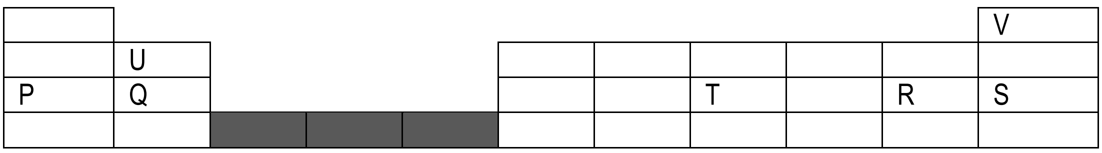
    (c) Calcium metal reacts with water to form calcium hydroxide solution and hydrogen gas. (2 marks)

36. Zinc reacts with dilute sulphuric acid to produce a gas which is colourless
    a) Identify the gas that was produced (1 mark)

    b) State two properties of this gas (1 mark)

    c) Write an equation for the reaction between zinc and dilute sulphuric acid (2 mks)

37. The figure below represents part of the periodic table. The letters do not represent actual symbols of the elements. Use it to answer the questions that follow.



a) What is the name for the elements that occupy the shaded region? (1 mark)

b) Select two elements which belong to the same group. (1 mark)

c) State the period to which element T belongs (1 mark)

d) Write down the electron pattern for the atom of element R (1 mark)

e) State the valence electrons in the atom of element S (1 mark)

f) Indicate on the grid the position of element Y which belongs to period 3 and has an ionic charge of 3+. (1 mk)

38. Naturally occurring silver consists of two isotopes

$^{107}_{47}Ag$ and $^{109}_{47}Ag$

They occur in equal numbers. Calculate the relative atomic mass of silver.

39. An element whose atomic number is 14 has three isotopes A, B and C. Use the information given in the table below to answer the questions that follow.

Isotopes	No of neutrons	% abundance
A	14	92.0
B	15	-
C	16.	3.0.

a) Determine the relative abundance of the isotope B (1 mark)

b) Calculate the relative atomic mass of the element above. (3 marks)

40. Write down the correct chemical equations for the word equations below and then balance them fully.

(a) Sodium oxide + water $\longrightarrow$ sodium hydroxide (2 marks)

(b) Copper + 0xygens $\longrightarrow$ copper (II) oxide (2 marks)

(c) Chlorine gas + Hydrogen gas $\longrightarrow$ Hydrogen chloride (2 marks)

(d) Carbon + Oxygen $\longrightarrow$ carbon (II) oxide (2 mks)

41. Write down the electron pattern for the following

i. Boron (5 electrons)

ii. Fluorine (9 electrons)

iii. Phosphorus (15 electrons)

iv. Hydrogen (1 electrons)
v. Neon (10 electrons)

42. Draw the electrons structure of the atom of an element represented as.(2 marks)

$^{40}_{20}X$

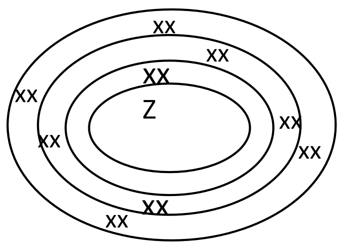
43. Write down the electron pattern for the ions of the atoms shown below.

i) $^{24}_{12}M$	iv) $^{27}_{13}Al$
ii) $^{32}_{16}S$	i) $^{35}_{17}Cl$

44. Calculate the relative atomic mass of an element Y whose isotopic composition is as follows: (3 marks)

$^{63}_{29}Y$ 71% and $^{56}_{29}Y$ 29%

45. An atom if an element Z with mass number 40 is represented below.



Determine:

a) The atomic number of Z. (1mk)

b) The number of electrons in Z (1 mk)

c) The number of neutrons in Z (1 mk)

d) The period of Z (1 mk)

e) The group of Z (1 mk)

46. The electron arrangement of ions X^{3+} and Y^{-2} are 2:8 and 2:8:8 respectively.
    a) Write the electron arrangement of elements X and Y (2 marks)

    b) Write the formula of the compound that would be formed between X and Y (1 mark)

47. An element Q can be represented as $^{40}_{20}Q$ (Q is not necessarily the symbol of the element). Identify its
    a) Mass number (½ mark)
    b) Atomic number (½ mark)
    c) Number of Neutrons (½ mark)
    d) The period and group in which element Q belongs. (1 mark)
    e) Draw the atomic structure of Q showing the electron arrangement. (1 mark)

48. Given that elements X has atomic number 16. State the:
    a) Number of electrons of atom X (1 mark)
    b) Electron arrangement of X (1 mark)
    c) The formula of the Ion of X. (1 mark)

49. a) An element Q has two Isotopes with relative abundance of 65% and 35%. If the mass number of the two isotopes is X and 31 respectively. Find the mass number represented by X given the R.A.M. of element Q is 30. (3 marks)

    (b)An element B consists of two Isotopes B – 35 and B – 37 in the ratio 3:1 respectively. Calculate the R.A.M. of B. (3 mks)

50. The table below gives elements represented by letters T – Y and their atomic number. (The letters do not represent the actual chemical symbols of the elements)

Elements	T	U	V	W	X	Y
Atomic numbers	12	13	14	15	16	17
Electron arrangements						

    a) Complete the table giving the electron arrangement of each of the element (3 marks)
    b) In which period of the periodic table do these elements belong. (1 mark)
    c) Write the formula of the compound formed between T and X. (1 mark)
    d) Write the formula of the carbonate of element U. (1 mark)

51. Element X,Y and Z have the following electronic configuration.

X – 2.2
Y – 2.8.2
Z – 2.8.8.2

a) To which group of the periodic table do element X,Y, and Z belong. (1 mark)

b) What is the valency of the elements in the group. (1 mark)

c) Write the formula of their:

i. Carbonates

ii. Nitrates of the above elements (3 marks)

52. The grid given below represents part of the periodic table. The letters are not actual symbols of the elements.

							V
	U					H	
G	Q	K	T			D	S

a) Write down the electronic configuration of elements Q (1 mark)

b) Write the symbol Ion of D (1 mark)

c) Write the formula of the compound formed when:

i. G combined with D.

ii. K combined with H (2 marks)

d) Element M has atomic number 15. Locate its position using (X) in the grid above. (1 mark)

53. Write the formula of the following compounds. (5 mks)

(i)Lithium chloride.

(ii)Zinc (11)chloride

(iv)Zinc (ii) sulphate

(iii)Calcium carbonate-

(v)Lead (11) Nitrate

54. The valencies of metals X, Y and Z are 1, 2 and 3 respectively. Write the formula of their:- (6 marks)

a) Hydroxides (OH)

b) Sulphates(SO_4^{2-})

55. An ion of phosphorous can be represented as $^{31}_{15}P^{3-}$

Draw a diagram to show the distribution of the electrons and the composition of the nucleus of the ion of phosphorous.

56. Elemement P,Q,R,and S have atomic numbers 2,10,12,and 17 respectively , select two elements that belong to the same group (2 marks)

CHAPTER TWO: CHEMICAL FAMILIES

Specific Objectives

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

a) identify alkali metals, alkaline-earth metals, halogens and noble gases in the periodic table and write their electron arrangement

b) state and explain trends in physical properties of alkali metals, alkaline-earth metals, halogens and noble gases

c) state and explain the trends in reactivity of the alkali metals, alkaline-earth metals and halogens

d) explain the similarities in formulae of compounds formed by alkali metals, alkaline-earth metals and halogens

e) state the uses of alkali metals, alkaline-earth metals, halogens and noble gases

f) explain the unreactive nature of the noble gases in terms of their electron arrangement

g) identify the elements in a given period and write their electron arrangement

h) state and explain the trends in physical properties of elements in a period

i) state and explain the trend in chemical behaviour of elements in a given period

INTRODUCTION

Elements in the same group have similar chemical properties because they have the same number of electrons in the outermost energy level. Elements in the same group belong to the same chemical family. There are four chemical families in the periodic table

1. Group I – alkali metals
2. group II – alkaline earth metals
3. group VII –halogens
4. Group VIII – noble gases

ALKALI METALS

These are elements in group I of the periodic table and include Lithium, sodium, potassium, rubidium, cesium and francium. They have one electron in their outermost energy level and are therefore monovalent

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

PHYSICAL PROPERTIES

- Have a metallic luster/shine when polished
- Good conductors of heat and electricity
- Are soft and can be easily cut with a knife
- Relatively low melting point and boiling point due their weak metallic bond as they have a bigger atomic radius
- Less dense than water

Group 1 element are usually stored under paraffin to prevent them from coming into contact and reacting with air and moisture.

Summary of physical properties

Element	**appearance**	**Ease of cutting**	**Melting point °C**	**Boiling point °C**	**Electrical conductivity**	**Atomic radius**	**ionic radius**	**1st ionization energy**
Lithium	Silvery white	Slightly hard	180	1330	Good	0.133	0.060	520
Sodium	Shiny grey	Easy	98	890	Good	0.157	0.095	496
Potassium	Shiny grey	Easy	64	774	good	0.203	0.133	419

Trends in physical properties down the group

- Alkali metals have relatively low meting point and boiling point. The melting point and boiling point decrease down the group due to decrease in strength of metallic bonding with increase in atomic size. **NB**; strength of metallic bonding is determined by the atomic radius the larger the atomic radius the weaker the metallic bond
- Ease of cutting increase down the group due to decrease in strength of metallic bond with increase in atomic size.
- Atomic masses increase down the group due to increase in number of protons and neutrons
- The atomic radius and ionic radius **increase down the group** due to increase in number of **occupied** energy levels
- The atomic radius is greater than ionic radius because the alkali metals react by losing one electron in the outermost energy level hence their ions have one less occupied energy levels
- The ionization energy decrease down the group due to increase in atomic radius which decreases the force of attraction between the outermost electron and the positive nucleus.
- The electrical conductivity increase down the group due to decrease in force of attraction between the nucleus and outer electron with increase in atomic radius hence the electrons are more delocalised

CHEMICAL PROPERTIES

Alkali metals react by losing outermost electrons to form a monovalent cation.

$$\mathbf{Li}_{(s)} \longrightarrow \mathbf{Li}^{+}_{(q)} + \mathbf{e}^{-} \quad 1^{st}\ ionization\ energy\ 520\ kJ/mol$$

$$\mathbf{Na}_{(s)} \longrightarrow \mathbf{Na}^{+}_{(g)} + \mathbf{e}^{-} \quad 1^{st}\ ionization\ energy\ 496\ kJ/mol$$

$$\mathbf{K}_{(s)} \longrightarrow \mathbf{K}^{+}_{(g)} + \mathbf{e}^{-} \quad 1^{st}\ ionization\ energy\ 419\ kJ/mol$$

The ionization energy is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms.
NB .***reaction is endothermic ΔH = +ve***

- The ionization energy **decrease down the group** due to increase in **atomic radius** which decreases the attraction between the outermost electron and the positive nucleus. hence potassium having the largest atomic radius amongst the first three alkali metals ,require least energy to remove the outermost electron.
- Therefore is the most reactive/most electropositive element among the first 3 group one elements

-Lithium, sodium and potassium react vigorously with both air and water and that is why they are **stored under paraffin**.
-When sodium is exposed it reacts with moisture in the air to form sodium hydroxide, which further reacts with CO_2 in the air to form sodium carbonate

$$2Na_{(s)} + H_2O_{(s)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

$$2NaOH_{(aq)} + CO_{2(g)} \longrightarrow Na_2CO_3.H_2O_{(s)}$$

i) Reaction of alkali metals with air /oxygen

-Sodium burns in air with a **yellow flame** to form a white solid which is a mixture of sodium oxide and sodium nitride, but sodium oxide is a major component because oxygen is more reactive than nitrogen.

$$4Na_{(S)} + O_{2\,(g)} \longrightarrow 2Na_2O_{(s)}$$

$$2Na_{(s)} + N_{2(g)} \longrightarrow 2Na_3N_{(s)}$$

-When water is added to the above solid alkaline solution is formed which is sodium hydroxide, when water is added to sodium nitride ammonia gas is evolved which has a characteristic pungent smell and turns red litmus turns blue.

$$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$

$$Na_3N_{(s)} + 3H_2O_{(l)} \longrightarrow 3NaOH_{(aq)} + NH_{3(g)}$$

-Sodium burns in air enriched with oxygen to from mainly sodium peroxide(yellow)

$$2Na_{(S)} + O_{2\,(g)} \longrightarrow Na_2O_{2(s)}$$

When water is added to the sodium peroxide gas that relights a glowing splint (oxygen) is evolved

$$2Na_2O_{2(s)} + 2\,H_2O_{(l)} \longrightarrow 4NaOH_{(aq)} + O_{2\,(g)}$$

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

$$4K_{(s)} + O_{2\,(g)} \longrightarrow 2K_2O_{(s)}$$

Lithium burns in air with a **red flame** forming lithium oxide

$$4Li_{(s)} + O_{2\,(g)} \longrightarrow 2L_2O_{(s)}$$

Nb; reactivity of group I metals increase down the group due to increase in atomic radius which decrease the force of attraction between the nucleus and outer electrons and also increase in shielding effect hence the electron is easily lost in larger atoms.

ii) Reaction with water.

Alkali metals react with water to form **metal hydroxide** and **hydrogen gas**. The reactivity of alkali metals increase down the group due to increase in atomic radii which increase the ease in which the atom loses the outermost electrons.

<u>Observations when alkali metals are placed in water</u>

Alkali metal	**Observations**	**Comparative speed/rate of the reaction**
Lithium	-Metal floats in water -effervescence//bubbles of colourless gas produced that extinguishes burning splint with a pop" sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Moderate
Sodium	-Metal floats in water, melts into silvery ball, darts on the surface of water , very rapid effervescence /bubbles of a , colourless gas produced that extinguishes a burning splint with a"pop" sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Very vigorous
Potassium	-Metal floats in water,-melts into a silvery ball/burst into a lilac flame -darts on the surface of water bubbles of a 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

<u>Explanation</u>

Alkali metals are less dense than water. They therefore float in water. the reaction produced hydrogen gas hence effervescence, the metal melts into a silvery ball because the reaction is exothermic,and hydrogen burns with a pop sound ,the hydrogen produced propels the metal hence darts on the surface of water, They react with water to form a strongly alkaline solution of their hydroxides and hence turns phenolphthalein indicator pink/turns red litmus paper blue. The rate of this reaction increase down the group. i.e. Potassium is more reactive than sodium .Sodium is more reactive than Lithium.

chemical equations

$$2Li_{(s)} + 2H_2O_{(l)} \longrightarrow 2LiOH_{(aq)} + H_{2(g)}$$

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

$$2K_{(s)} + 2H_2O_{(l)} \longrightarrow 2KOH_{(aq)} + H_{2(g)}$$

Reactivity increase down the group

NB: potassium bursts into a lilac flame when reacted with water. The reaction ios highly exothermic such that the hydrogen gas produced ignites. It burns with lilac flame due to presence of potassium vapour.

iii) Reaction with chlorine

Alkali metal reacts with chlorine to form the corresponding metal chlorides, reactivity increase down the group

When hot sodium metal is lowered into a chlorine gas, it bursts into a yellow flame forming white fumes of sodium chloride.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

$$2Li_{(s)} + Cl_{2(g)} \longrightarrow 2LiCl_{(s)}$$

$$2K_{(s)} + Cl_{2(g)} \longrightarrow 2KCl_{(s)}$$

Similarity of ions and formulae of the compounds formed between an alkali metal ion and hydroxide.

Akali metal ion	**OH^-**	**SO_4^{2-}**	**PO_4^{3-}**	**ClO_4^-**	**HCO_3^-**	**Nitride N^{3-}**
Li+	LiOH			$LiClO_4$		
Na+		Na_2SO_4				
K+			K_3PO_4			

USES OF ALKALI METALS AND THEIR COMPOUNDS

1. Sodium is used in making of sodium cyanide is used in extraction of gold
2. Lithium is used in manufacture of high special strength glasses and ceramics
3. Lithium compounds are used in the manufacture of dry cells for use in mobile phones, laptops etc
4. Sodium vapour is used to produce yellow glow in streetlights.
5. Sodium chloride is used as a food additive.
6. Sodium hydroxide is used in manufacture of soap.
7. Sodium is used as a reducing agent in extraction of titanium from titanium (IV) chloride.
8. A molten mixture of sodium and potassium are used as a coolant in nuclear reactors.

ALKALINE EARTH METALS

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

Group two elements are called alkaline earth metals because;

- They are abundant in the earth's crust
- They dissolve in water to form alkaline solutions.

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

PHYSICAL PROPERTIES OF ALKALINE EARTH METALS

Alkaline earth metals are better conductors/have higher electrical conductivity than alkali metals because they have more delocalized electrons than alkali metals.
Alkaline earth metals have relatively stronger metallic bonding than alkali metals because they have smaller atomic radii and more valence electrons.

- good conductors of heat and electricity
- Hard and cannot be easily cut with a knife
- Have a dull appearance due to coating of their oxides but when freshly cut they have a metallic luster
- Have higher melting point and boiling point than alkali metals due their stronger metallic bonding
- Are denser than water that is why they don't float on water/ They sink in water.
- Are malleable (can be hammered into sheets)
- Are ductile (can be drawn into wires)

Trends in physical properties of group 2 elements

Element	symbol	Electronic configuration	Atomic radius	Melting Point	Boiling point	Ionic radius	First ionization energy	second ionization energy
Beryllium	Be	2.2	0.089	1280	2450	0.031	900	1800
magnesium	Mg	2.8.2	0.136	650	1110	0.065	736	1450
Calcium	Ca	2.8.8.2	0.174	850	1140	0.099	590	1150

Trends in physical properties

- Atomic radius and ionic radius increase down the group due to increase in number of occupied energy levels.
- Atomic masses increase down the group due to increase in number of protons and neutrons
- The atomic radius is greater than ionic radius because the atom reacts by losing two electrons in the outermost energy level and therefore the ion has one less occupied energy level than the atom.
- Ionization energy decrease down the group due to decrease in the force of attraction between the positive nuclei and the outermost electron with increase in atomic radius.
- The second ionization energy is greater than the first ionization energy because when the first electron is lost the overall positive charge attracts the remaining electron more firmly.

Alkaline earth metals have a higher boiling points and melting points than alkali metals because alkaline earth metals have more delocalized electrons hence stronger metallic bonding than alkali metals.

Alkali earth metals react by losing two electrons to become stable and therefore have valence of 2 and oxidation state of +2

$$Mg \longrightarrow Mg^{+} + e \left(1^{st}\ I.E = 736 \right)$$

$$Mg^{+} \longrightarrow Mg^{2+} + e\left(2^{nd}\ I.E = 1450 \right)$$

CHEMICAL PROPERTIES OF ALKALINE EARTH METALS

i) Burning in air

Magnesium burns in air with a **brilliant white flame** to form a white solid which is a mixture of MgO and Mg_3N_2

$$2Mg_{(S)} + O_{2\ (g)} \longrightarrow 2MgO_{(s)}$$

$$3Mg_{(s)} + N_{2(g)} \longrightarrow Mg_3N_{2(s)}$$

Calcium also burns in air with a **red flame** forming a white solid. Which is also a mixture of Calcium Oxide and Calcium Nitride.

$$2Ca_{(S)} + O_{2\ (g)} \longrightarrow 2CaO_{(s)}$$

$$3Ca_{(s)} + N_{2(g)} \longrightarrow Ca_3N_{2(s)}$$

Reaction with steam (refer to notes on water and hydrogen

ii) Reaction with dilute acids.

Alkaline earth metals react with dilute acid to form the corresponding salt and hydrogen gas. The reaction is accompanied by effervescence and generates heat.

$$Mg_{(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_{2(g)}$$

$$Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$$

$$Ca_{(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + H_{2(g)}$$

NB – the reaction between calcium and dilute sulphuric acid starts and then stops due to formation of **insoluble coat** of $CaSO_4$ **which coats** the metal preventing further reaction between the acid and the metal

Alkaline earth metals react by losing two electrons and therefore have valency 2 (divalent) they form compounds with similar formulae.

The table below shows some compounds of some alkaline earth metals

Radicle	**Beryllium**	**Magnesium**	**Calcium**	**Barium**
Hydroxide	$Be(OH)_2$	$Mg(OH)_2$	$Ca(OH)_2$	$Ba(OH)_2$
Oxide				
Sulphide				
Chloride				
Carbonate				
Nitrate				
Sulphate				
Sulphite				
Hydrogen carbonate				
Hydrogen sulphate				

Uses of alkaline earth metals and their compounds

1. Magnesium is used to make Magnesium hydroxide is used in manufacture of anti-acid medicine
2. Magnesium –aluminium alloy is used to make aeroplane parts
3. Calcium is used to make Hydrated $CaSO_4$ is used in hospitals to set fractured bones
4. $CaCO_3$ is used in extraction of iron to remove impurities and manufacture of cement
5. $CaCO_3$ is used in solvay process to make sodium carbonate
6. CaO is used to raise the soil pH for agricultural purposes
7. $Ca(OH)_2$ is used to prepare ammonia gas in the laboratory
8. Calcium nitrate is used as a nitrogenous fertilizer
9. MgO is used in lining of furnaces
10. Barium sulphate is used in diagnosis of ulcers
11. Barium nitrate is used to produce green flame in fireworks
12. $CaCO_3$ is mixed with oil to make putty

USE OF IONIZATION ENERGIES TO DETERMINE THE GROUP OF AN ELEMENT.

In such questions whereby you are required to ***use of ionization energies to determine the group of an element*** you look for abnormally big difference between two ionizations energies, that gives indication that the electron is being removed from a stable ion e.g study ionization energies in kilojoules per mole and answer the questions below

Element	1st	2nd	3rd	4th	5th	6th
A	1590	2780	4700	6500	8100	12500
B	1010	1900	4900	5000	6300	7300
C	940	4800	6300	9100	12000	16500
D	1680	2010	3400	10900	12400	15900

Identify the group to which each element belongs to

- ✓ Element A –the difference between the 5th and 6th ionization energy is very high ,that means the sixth electron is being removed from a stable atom hence A is in group V
- ✓ Element B is in group II because there is a big difference between 2nd and 3rd, meaning the 3rd electron is being removed from a stable atom.
- ✓ Element C is in group I and D is in group III

HALOGENS

These are elements in group VII of the periodic table and include;

Halogen	symbol	Atomic number	Electron arrangement
Fluorine	F	9	2.7
Chlorine	Cl	17	2.8.7
Bromine	Br	35	2.8.18.7
Iodine	I	53	2.8.18.18.7

Appearance of halogens

Halogens are coloured and some halogens show different colours depending on the state.

Some adjectives like yellowish ,reddish ,brownish should not be used to describe colour but simply use yellow, red, brown e.t.c

Fluorine – yellow gas

Chlorine – green gas /pale green gas

Bromine water is **yellow**,

bromine liquid is *red* and bromine gas is brown

Iodine – black solid , iodine solution is solution and a purple gas/vapour

The colour of halogens depends on the state as shown above

PHYSICAL PROPERTIES OF HALOGENS

halogen	Symbol	Atomic number	Atomic radius	Ionic radius	Melting point	Boiling point	Appearance at room temp.	Electron affinity
flourine	F	9	0.064	0.136	-238	-188	Yellow gas	-322
Chlorine	Cl	17	0.099	0.181	-101	-35	Green gas	-349
Bromine	Br	35	0.114	0.195	-7	59	red liquid	-325
Iodine	I	53	0.133	0.216	114	184	Black solid	-295

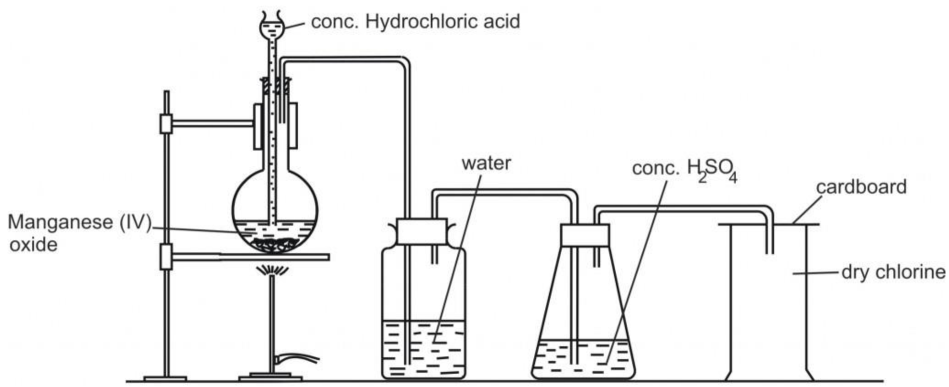
Trends in physical properties

- The atomic radius and ionic radius **increase down the group** due to increase in number of occupied energy levels.
- Atomic masses increase down the group due to increase in number of protons and neutrons
- The ionic radius is greater than atomic radius because halogens react by gaining electrons which leads to increase in repulsion force hence increase in the ionic radius
- The melting point and boiling point increase down the group due to increase in intermolecular forces of attraction between molecules with increase in size of the molecules.

- Fluorine and chlorine are gases at room temperature, bromine is a liquid and iodine is a solid due to increase in intermolecular forces of attraction between molecules with increase in size of the molecules .
- The electron affinity of halogens decrease down the group due to increase in atomic radius and therefore reactivity decreases down the group.

LABORATORY PREPARATION OF CHLORINE GAS

It is prepared in the lab by heating concentrated HCl with Manganese (IV) oxide.



$$4HCl_{(aq)} + MnO_{2(s)} \longrightarrow MnCl_{2(aq)} + 2H_2O_{(l)} + Cl_{2(g)}$$

MnO_2 is an oxidizing agent we can also use $KMnO_4$ but in this case there is no heating because $KMnO_4$ is a stronger oxidizing agent

$$2KMnO_{4(s)} + 16HCl_{(aq)} \longrightarrow 2KCl_{(aq)} + 2\ MnCl_{2(aq)} + 5Cl_{2(q)} + 8H_2O_{(l)}$$

Chlorine can also be prepared in the lab by reacting bleaching powder calcium hypochlorite with dilute acids

$$2CaOCl_{(aq)} + 4HNO_{3(aq)} \longrightarrow 2Ca\left(NO_3\right)_{2(aq)} + 2H_2O_{(l)} + Cl_{2(g)}$$

Chlorine is also prepared by reacting lead (IV) oxide with concentrated HCl(heating is required)

$$4HCl_{(aq)} + PbO_{2(s)} \longrightarrow PbCl_{2(s)} + 2H_2O_{(l)} + Cl_{2(g)}$$

Test for chlorine gas

It will turn moist blue litmus paper red then white

CHEMICAL PROPERTIES OF HALOGENS

i) Reaction between chlorine and water

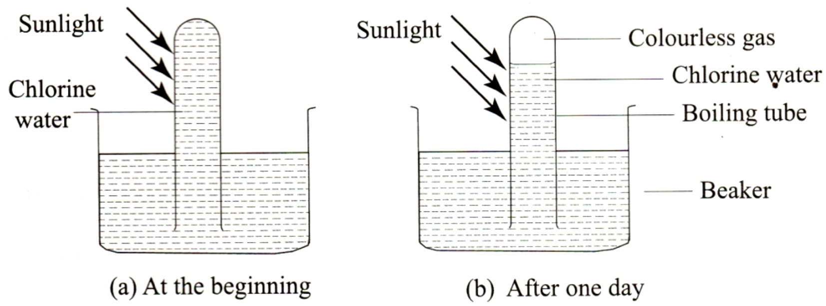
Chlorine gas dissolves in water to form a yellow solution called chlorine water, which is a mixture of hydrochloric (HCl) acid and choric (I) acid (Hypochlorous acid). (HOCl)

$$Cl_{2(g)} + H_2O_{(l)} \xrightarrow{sunlight} HCl_{(aq)} + HOCl_{(aq)}$$

Chlorine water turns blue litmus red then white

When chlorine water is left in sunlight, it loses its yellow colour and bubbles of a colourless gas are formed

This happens because HOCl is decomposed by sunlight to form HCl and oxygen.



$$2HOCl_{(aq)} \longrightarrow 2HCl_{(aq)} + O_{2\,(g)}$$

ii) Bleaching action of chlorine

Chlorine dissolves in water to form HCl and HOCl

In presence of sunlight HOCl decomposes to form HCl and Oxygen

Oxygen combines with the dye decolourising it, therefore chlorine bleaches through oxidation

$$Cl_{2(g)} + H_2O_{(l)} \longrightarrow HCl_{(aq)} + HOCl_{(aq)}$$

$$2HOCl_{(aq)} + dye \longrightarrow 2HCl_{(aq)} + \underbrace{(O + dye)}_{Colourles}$$

The bleaching aaction of chlorine is permanent.

iii) Reaction of chlorine with alkalis

a)with cold dilute NaOH

$$2NaOH_{(aq)} + Cl_{2(g)} \longrightarrow NaCl_{(aq)} + NaClO_{(aq)} + H_2O_{(l)}$$

$$2KOH_{(aq)} + Cl_{2(g)} \longrightarrow KCl_{(aq)} + KClO_{(aq)} + H_2O_{(l)}$$

b)with hot concentrated NaOH

$$6NaOH_{(aq)} + 3Cl_{2(g)} \longrightarrow NaClO_{3(aq)} + 5NaCl_{(aq)} + 3H_2O_{(l)}$$

$$6kOH_{(aq)} + 3Cl_{2(g)} \longrightarrow KClO_{3(aq)} + 5KCl_{(aq)} + 3H_2O_{(l)}$$

iv)Displacement reactions

When chlorine is bubbled into a solution of potassium bromide. The green colour of chlorine gas disappears, the solution turns from colourless to yellow

$$Cl_{2(aq)} + 2KBr_{(aq)} \longrightarrow 2KCl_{(aq)} + Br_{2(l)}$$

This is due to the formation of bromine liquid which is orange. It is a displacement reaction since chlorine is more reactive than bromine ,it will displace it from its compounds.

Similarly when chlorine is bubbled into a solution of potassium iodide, the solution turns brown and eventually a black solid is formed.

$$Cl_{2(g)} + 2KI_{(aq)} \longrightarrow 2KCl_{(aq)} + I_{2(aq)}$$

v)Reaction with non-metals

-Reaction with hydrogen gas

Hydrogen burns in atmosphere of chlorine forming hydrogen chloride

$$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(aq)}$$

That is how hydrochloric acid is formed in industries

-Reaction with iodine

When chlorine is reacted with iodine it form iodine chloride (amber coloured liquid)

$$Cl_{2(g)} + I_{2(s)} \longrightarrow 2ICl_{(l)}$$

If excess chlorine is used iodine trichloride is formed (yellow solid)

$$Cl_{2(g)} + I_{2(s)} \longrightarrow ICl_{3(s)}$$

***-Reaction with* Phosphorous**

Phosphorous burns with chlorine to form phosphorous (V) chloride and phosphorous (III) chloride

$$2P_{(s)} + 5Cl_{2(g)} \longrightarrow 2PCl_{5(g)}$$

$$2P_{(s)} + 3Cl_{2(g)} \longrightarrow 2PCl_{3(g)}$$

vi) Reaction with metals

Chlorine reacts with metals to form metal chlorides

Reacting chlorine with iron metal

A stream of dry chlorine is passed over heated iron wool as shown in the diagram below.iron glows red and reacts with chlorine forming brown crystals of iron (III) chloride. Heating is stopped when iron glows red ,this is because the reaction is exothermic and the heat generated by the reaction is enough to sustain the reaction temperature.

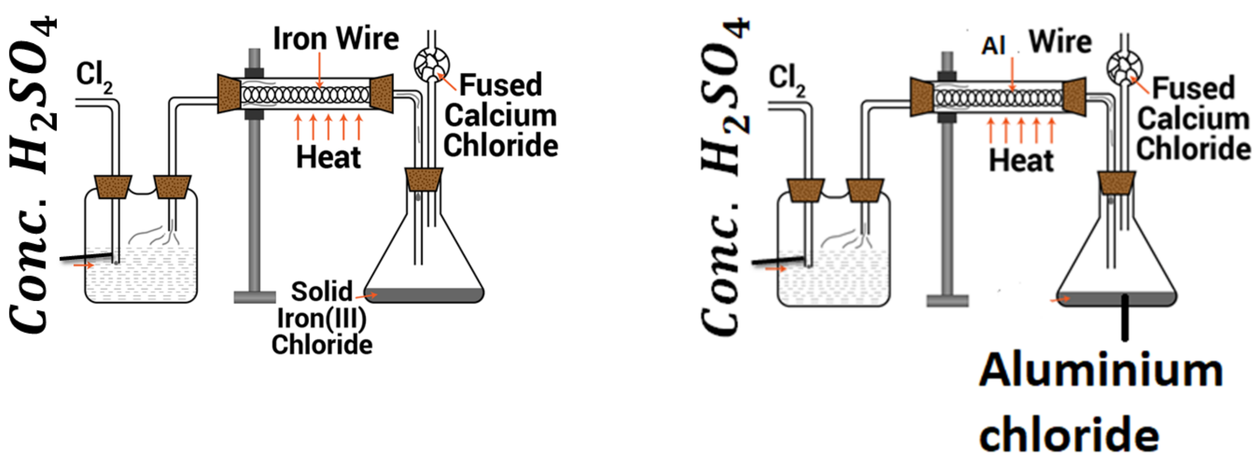
$$2Fe_{(s)} + 3Cl_{2(g)} \longrightarrow 2FeCl_{3(s)}$$

Iron (III) chloride appears in the form of a brown sublimate in the conical flask. This is because the heat from the combustion tube prevents it from forming there. The excess chlorine is fed to the fume chamber since it is poisonous or it can be absorbed by bubbling it into a concentrated NaOH solution.The$CaCl_2$ prevents the atmospheric moisture from entering the apparatus. This is because $FeCl_3$ is deliquescent. Instead of $CaCl_2$ you can also use CaO which has the added advantage of reacting with excess Cl_2 ,

$AlCl_3$ can be prepared using the same method but will appear as a white sublimate.

$$2Al_{(s)} + 3Cl_{2(s)} \longrightarrow 2AlCl_{3(s)}$$

Reaction of chlorine gas with Iron and Aluminium metals



Sodium metal burns in atmosphere of chlorine forming sodium chloride

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

Similarly halogens burns in heated zinc as shown in the equations below.

$$Zn_{(s)} + Cl_{2(g)} \longrightarrow ZnCl_{2(s)}$$

$$Zn_{(s)} + I_{2(s)} \longrightarrow ZnI_{2(s)}$$

$$Zn_{(s)} + Br_{2(g)} \longrightarrow ZnBr_{2(s)}$$

Magnesium also reacts with chlorine

$$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$$

vii)Chlorine as an oxidizing agent

When chlorine gas is bubbled through Iron (II) Sulphate/ Iron (II) Chloride solution the solution turns from green to yellow, this is because the green Fe^{2+} are oxidized to yellow Fe^{3+}

$$2FeCl_{2(s)} + Cl_{2(s)} \longrightarrow 2FeCl_{3(s)}$$

Uses of halogens

1. Fluorine is used in the manufacture of toothpaste and plastics (polytetraflouride)
2. Fluorine is used to manufacture of chloroflourocarbons (CFC_s) used as refrigerants
3. Chlorine is used as a bleaching agent in paper and textile industry
4. Chlorine is used in water treatment plants
5. Chlorine is used in manufacture of antiseptics e.g dettol
6. Chlorine is used in manufacture of PVC polyvinylchloride used to make plastic pipes
7. Chlorine is used in manufacture of hydrochloric acid
8. Bromine is used in lab to test for unsaturation
9. Bromine is uded to make flame retardants
10. Iodine is used to make iodised table salt
11. Iodine is used to make antiseptics e.g tincture of iodine (Mixture of Ethanol and Iodine)

NOBLE GASES

These are elements in group VIII of the periodic and include .helium, neon, argon, krypton and xenon. They have a stable electron arrangement and therefore do not react under normal conditions. Therefore they are the least reactive elements.

They were previously known as rare gases (inert gases) but these names proved to be misnomer. This is because argon at 0.93% is more abundant than carbon (IV) oxide at 0.03% .some of the noble gases can be reacted under extreme conditions e.g.

$$Xe_{(s)} + 2F_{2(g)} \longrightarrow XeF_{4(s)}$$

But generally noble gases are inert as they have a stable configuration and exist as monoatomic gases (single) atoms. They have valence zero

Trends in physical properties of noble gases

Element	Symbol	Appearance At room temp.	Atomic number	Atomic radius	1st ionizatuion energy	Melting point	Boiling point
Helium	He	Colourless gas	2	0.128	2372	-272	-269
Neon	Ne	Colourless gas	10	0.160	2080	-249	-246
Argon	Ar	Colourless gas	18	0.192	1520	-189	-186
Krypton	Kr	Colourless gas	36	0.217	1350	-157	-152

physical properties of noble gases

- Colourless gases at room temperature
- Atomic radii increase down the group due to the increase in number of occupied energy levels; this explains why the1stionizatuion energy also decreases down the group.
- Atomic masses increase down the group due to increase in number of protons and neutrons
- Noble gases have low Melting point and Boiling point due to **weak van der waals forces** of attraction between the atoms, the strength of inter atomic forces of attraction increase with increase in atomic size and that is why melting and boiling point **increase down the group**.
- Noble gases have the highest ionization energy because they have a stable electron arrangement .

TRENDS IN CHEMICAL PROPERTIES OF NOBLE GASES

Noble gases have a stable configuration as their outermost energy level is filled up, hence do no react under normal conditions.

However as you move down the group ,atoms of noble gases with large atomic radius like krypton and xenon are able to lose electrons as they are very far from the nucleus and they are weakly attracted by the positive nucleus. That's why some compounds of some noble gases are known to exist.e.g **Xenon hexafluoride** ($XeFe_6$)

USES OF NOBLE GASES

Neon is used in advertisement and disco lights

- Argon and krypton is used to fill electric bulbs fluorescent tubes to prevent oxidation of the filament
- Helium and argon is used to provide inert atmosphere during welding
- Helium is used in filling airships to increase their buoyancy
- A mixture of helium and oxygen is used in deep sea diving and mountaineering
- Helium is used in filling weather balloons ,(it has replaced hydrogen which is more reactive)
- Liquid Helium is used as a coolant in nuclear reactors
- Neon is used as a refrigerant
- Neon is used in advertisement and disco light as it glows when electric current is passed through it .
- Neon is used in television CRO to bring about colour
- Xenon is used in cameras to produce flash light
- Xenon is used in disco lights

REVISION QUESTIONS CHEMICAL FAMILIES

1. Write a balanced chemical equation for the following reactions.
   i. Calcium + water (1 mark)

   ii. Sodium + chlorine gas (1 mark)

   iii. Magnesium + Hydrochloric acid (1 mark)

2. A science student discovered a new element in the school compound and named it "pelenium". It was a solid she could easily cut with a knife. When she put it into water, it reacted vigorously and caught fire forming an alkaline solution. The symbol of peleunium is is 'Pe'
   i. How did she store it? (1 mark)

   ii. Was it a metal or a non-metal? Give a reason (2 marks)

   iii. Which chemical family does "Pe" belongs? (1 mark)

   iv. What is the valency of the element? (1mk)

   v. Write the formulae for its;
      a) Chloride (1 mark)

      b) Carbonate (1 mark)

3. A piece of sodium metal is put on a basin of water.

   a) State the observations (2 marks)

   b) Write a word equation for the reaction between sodium metal and water.(1 mark)

   c) State the effects of the resulting solution on blue and red litmus paper.(1 mark)

4. A certain element Z has electron arrangement as 2.8.3
   a) What is the formula of the ion (1 mark)

   b) What is the oxidation number of its ion (1 mark)

   c) When alkali metals are freshly cut, they look shiny and silvery. However after a short while they tarnish (become dull). Explain (1 mark)

5. The following table shows the physical properties of the first 3 alkali metals.

Element	**Atomic number**	**Atomic radius**	**Ionic radius**	**Meilting point**	**Boiling point**	**Ionization energy**
Lithium	3	0.133	0.060	180	1330	520
Sodium	11	0.157	0.095	98	890	496
potassium	19	0.203	0.133	64	774	419

   a) Explain the trend in: (3 mks)
      i. Melting and boiling points.

      ii. Atomic and ionic radii

      iii. Ionization energy

   b) Comment on the thermal and electrical conductivity of alkali metals.(1 mark)

   c) Explain why the ionic radius is smaller than the atomic radius (2mks)

6. A student placed a small piece of sodium metal in a trough of water.
   a) State and explain the observation made (1 mark)

   b) Red and blue litmus papers were added to the resulting solution. State and explain observation made. (1 mark)

   c) Write the balanced equation for the reaction that take place. (1 mark)

   d) Give one use of sodium metal (1 mark)

7. (a)Magnesium when burnt in air (oxygen) it forms an oxide, when the oxide was dissolved in water and the resulting solution was tested using red litmus paper.
   i. Write a balanced chemical equation when magnesium is burnt in oxygen. (1 mark)

ii. State the observation in (a(i) above. (1 mark)

iii. Write a chemical equation when the resulting product (oxide) was dissolved in water (1 mark)

iv. What was the effecting of the resulting solution formed in a(iii) above on litmus papers.

8. (b)Study the following equation and the amount of energy involved

$Mg_{(g)} \longrightarrow Mg^{+}_{(g)} + e$ (1[st] ionization energy = 736kJmol)

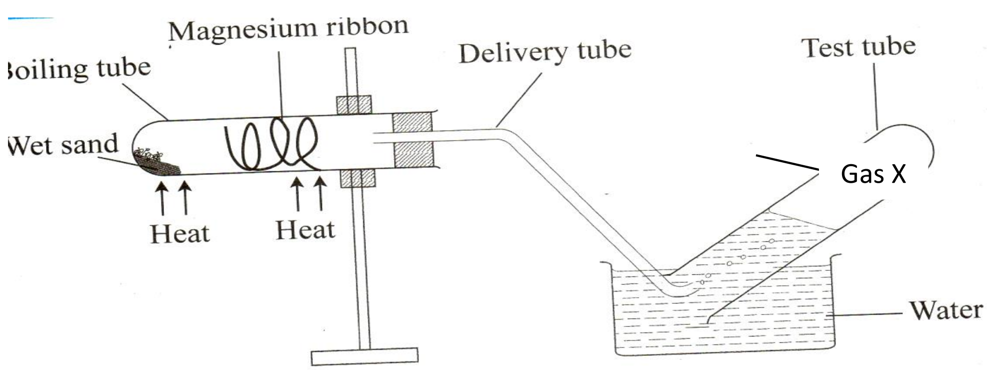
$Mg^{+}_{(s)} \longrightarrow Mg^{2+}_{(g)} + e$ (2[nd] ionization energy = 1150KJlmol)

i. Comment on the 1[st] and 2[nd] ionization energy. (1 mark)

ii. Explain your answer basing on the comment in the b(i) above. (2 marks)

iii. Give one use of magnesium (1 mark)

9. The following set-up was used to collect a certain gas (x) in the laboratory by students in form 2N of Gatondo Girls High school. Study it and answer the questions.



a) Name the gas X (1 mark)

b) How can you test the gas X in the laboratory? (1 mark)

c) Give the colour of the solid form at the end of the experiment. (1 mark)

d) Write a balanced equation to justify your answer in c above. (1 mark)

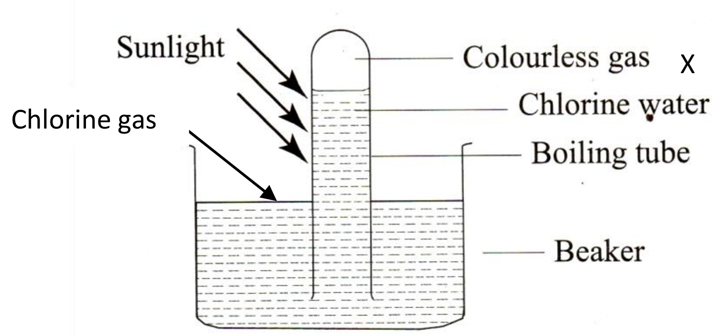
e) How did Magnesium burn in steam? (1 mark)

10. Write a chemical equation for the reaction between magnesium with: -
    a) Oxygen (1 mark)

    b) Chlorine (1 mark)

    c) Sulphuric (VI) acid (2 marks)

11. The following set-up was carried out by a student in form 2S in the open air in the presence of sunlight. The purpose of the experiment was to investigate a certain property of Halogen. Study it and answer the questions.



a) Which property of Halogens was being investigated? (1 mk)

b) Name gas which was colourless.

c) Chlorine water was yellow. Which substance is responsible for the yellow colour? (1 mk)

d) In presence of light, the chlorine water is decolourised. Explain. (1 mk)

e) Using equations, explain how the gas X is produced at the end of the experiment.
i. (2 mks)

f) Comment on effect of chlorine water on red and blue litmus paper. (1 mk)

g) Give one use of chlorine.

12. (a)Explain why group 8(noble gas) are unreactive gases (1 mark)

(b)Melting points and boiling points of Noble gas are relatively low. Explain. (1 mark)

13. Explain the trends in the following properties down the alkali metals
    a) Melting point and boiling point
    b) Ionization energy
    c) Atomic radius
    d) Ionic radius

14. The table below shows the trend in ionization energy for elements M, N, L .Use it to answer the questions that follow

Element	First Ionization energy(kJ)
M	494
N	519
L	418

a) Which elements have the smallest atomic radius? Give a reason.

b) Which element has the lowest melting point and boiling points?

c) Which elements would be a better conductor of both heat and electricity? Give a reason.

d) Which is the most reactive element

15. The table below shows the atomic and ionic radii of some elements in a group of a periodic table. The letters used do not represent the actual symbols of the elements.

Element	Atomic radii (nm)	Ionic radii (nm)
W	0.112	0.031
X	0.160	0.065
Y	0.197	0.099
Z	0.215	0.113

a) Explain why ionic radius is less than the atomic radius, for each of the elements. (1 mark)

b) Explain the variation in atomic radii down the group (1mark)

c) Identify the most reactive element (1 mark)

16. Study the information below and answer the questions that follow

Element	Atomic radii(nm)	Ionic radii(nm)
A	0.112	0.031
B	0.160	0.065
C	0.197	0.099
D	0.215	0.113

a) Would the above elements form a metallic or non metallic group? Explain (2mk)

b) Which is the most reactive element, explain (2mks)

17. The table below gives the ionization energy for group one elements

Element	I	II	III	IV
Ionization energy(KJ/mol-1)	94	418	519	576

Arrange the elements in order of reactivity starting with the most reactive

18. The table summarizes properties of alkaline earth metals.

Element	Atomic radius	Melting point	Boiling point	Atomic radius	1st ionization energy	2nd ionization energy
Be	4	1280	2770	0.11	900	1800
Mg	12	650	1110	0.16	740	1450
Ca	20	838	1440	0.20	590	1150
Sr	38	768	1380	0.21	550	1060

a) Account for the trends in each of the following;
Trend in atomic radius

Trend in 1st ionization energies

b) Account for the difference in first and second ionization energies

c) State two applications of alkaline earth metals in the health sector.

19. Study the table below and answer the questions that follow

															I	J	
	H														M	N	Q
	L																
K	O			P													
					R												
												S					

a) Explain how the reaction between chlorine and each of the following elements would compare

i. H and L

ii. K and O

b) How does the reactivity of I and M compare, explain?

c) Element X in period 3 and forms ions with formular X^{3-} place it in the grid above

d) Select the element that is least reactive, explain

e) Write the electronic arrangements of the ions of elements H and N

f) What name is given to elements in which P and R belong

20. Study the table below and answer the questions that follow

Element	Atomic number	Melting point	Formula of oxide
A	11	98	A2O
B	12	65	BO
C	13	660	C2O3
D	14	1410	DO2
E	16	113/119	EO2
F	17	-101	F2O7
G	18	-189	-

a. Which of elements are likely to have no chemical properties?

b. Why does element E have two melting points?

c. Explain why melting point of D is higher than that of F.

d. Which of the oxides can dissolve in both acids and alkalis, explain?

e. How does reactivity of B and C compare? give an explanation

f. Determine the oxidation states of F in its oxides

21. Study the table below and answer the questions that follow

A	B											C	D		E	F	G
															M	N	Q
K	H																

a. What name is given to group of elements in which :

A and K belong

F and N belong

B and H belong

b. Which letters represents the list reactive elements

c. On the grid indicate with a tick, the position of element P that is in the third period and would form P^{3-}

d. Write the formula of the compound formed when the following elements react:

I. A and E

II. B and M

III. K and E

IV. C and F

V. D and N

e. Select one element from the grid which;

I. Forms a monovalent cation

II. Forms a divalent cation

III. Forms a trivalent cation

IV. Forms a monovalent anion

V. Forms a divalent anion

VI. Forms an acidic oxide

VII. Form a basic oxide

VIII. Form a soluble basic oxide(alkali)

IX. Is likely to form a covalent bond

X. Forms an oxide that can react with both NaOH and HCl

22. The grid below represents part of the periodic table. the letters do not represent the actual symbols of the elements.

D								
							T	
	L			Q	I	Z		
G	R		M				E	

a) Select the most reactive non-metal

b) Write the formula of the compound formed when D and Z react

c) Select the element that form and ion of charge +2

d) Which element has the least ionization energy

e) Which element is the strongest reducing agent?

f) Which element has the highest electron affinity

g) Suggest the pH of aqueous solution of chloride of G,and Q

h) Compare the atomic radius of D and G

i) Compare the ionic radius of L and R

j) Compare the atomic mass of T and E

k) Compare the melting point and boiling points of T and E

l) The configuration of ion of W^- is 2.8.8, place it on the grid above

m) Compare the atomic radius and ionic radius of G

n) Compare the atomic radius and ionic radius of T

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

		Ionization Energy/Mole	
Element	Electronic configuration	1st ionization energy	2nd ionization energy
A	2.2	900	1800
B	2.8.2	736	1450
C	2.8.8.2	590	1150

a) What chemical family do the elements A, B and C belong? (1 mark)

b) Write the formula and draw the electronic structure of an ion of B(2 marks)

i. Formula

ii. Electronic structure

c) I). What is ionization energy? (1 mark)

II). Explain the following:

1. The 1st ionization energy is lower than the second ionization energy. (2mk)

2. The 1st ionization energy of B is higher than that of C.(2mk)

d) Write a chemical equation for the reaction of element B with:

1. Air

2. Chlorine gas

3. Steam (water vapour)

24. Study the information in the table below and answer the questions that follow (the letters do not represent the actual symbol of the substances)

Substance	Melting Point (ºC)	Boiling Point (ºC)	Solubility in water	Density at room temperature g/cm3
H	-117	78.5	Very soluble	0.8
J	-78	-33	Very soluble	0.77 x 10-3
K	-23	77	Insoluble	1.6
L	-219	-183	Slightly soluble	1.33 x 10-3

a) Which substance would dissolve in water and could be separated from the solution by fractional distillation? Give a reason (2 marks)

b) Which substances is a liquid at room temperature and when mixed with water two layers would be formed? Explain (2 marks)

c) Which letter represents a substance that is gas at room temperature and which can be collected

i. Over water? Explain (2 marks)

ii. By downward displacement of air? (Density of air is 1.29×10^{-3}g/cm^3 at room temperature). Explain (2 marks)

25. The grid below represents part of the periodic table. The letters do not represent the actual symbols.

								A
B			X	G	P	Z	E	V
	J		I	L	R	T		
D	G						M	

a) Select the most reactive
Non-metal. (1 mark)

Metal. (1mk)

b) Write the formula of the compound consisting of (10mksmk)

I.	.D and Z only.	II.	X and Z only
III.	Carbonate of J	IV.	Oxide of B
V.	Hydrogen compound of G	VI.	chloride of X
VII.	Sodium compound of E	VIII.	Nitrate of B
IX.	sulphate of D	X.	Aluminium compound of Z

c) Select an element that can form an ion of charge (5mk)

i. +1
ii. -1
iii. +2
iv. +3
v. -3

d) Which element has the least ionization energy? Explain (2 marks)

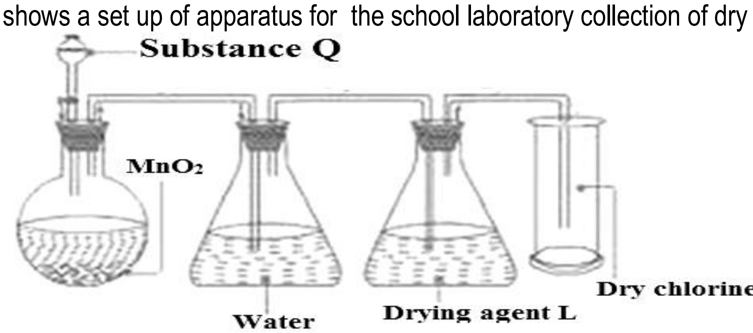
e) To which chemical family do the following elements belong? (3mk)

a. J
b. E
c. B

f) When a piece of element G is placed in cold water, it sinks to the bottom and effervescence of a colourless gas that burns explosively is produced. Use a simple diagram to illustrate how this gas can be collected during this experiment. (3 marks)

g) An element K has relative atomic mass of 40.2.It has two isotopes of masses 39 and 42. Calculate the relative abundance of each isotope. (3 marks)

26. The diagram below shows a set up of apparatus for the school laboratory collection of dry chlorine gas.



a) Name (2 marks)
   I. Substance Q

   II. Suitable drying agent L

b) State a missing condition for the reaction to take place faster. (1 mark)

c) Moist red and blue litmus papers were dipped into the chlorine gas from the above set up. State and explain the observations made. (2 marks)

d) Write the equation for the reaction taking place in the conical flask (1 mark)

e) Name two other substances that can be used in place of MnO_2 (2 marks

f) State three uses of chlorine (3 marks)

27. The number of protons, neutrons and electrons in atoms **A** to **F** are given in the table below the letters do not represent the actual symbol of the elements: **-**

Atoms	Protons	Neutrons	Electrons
A	3	4	2
B	9	10	10
C	12	12	12
D	17	18	17
E	17	20	17
F	18	22	18

Choose from the table the letters that represent:

i. An atom of a metal ....
ii. A neutral atom of a non-metal ....
iii. An atom of a noble gas …...
iv. A pair of isotopes ....
v. A cation ....

28. Use information in table below to answer questions that follow.

Ion	Electronic configuration of ion	Atomic Number	Electronic Configuration of Atom	Group	period
P^{2+}	2, 8, ,8				
Q^{-}	2, 8				
R^{-}	2, 8, 18, 8				
S^{+}	2, 8, 8				
T^{3+}	2, 8				
U^{2+}	2, 8, 18				
V^{2+}	2, 8				
W^{+}	2, 8				
X^{+}	2				
Y^{-}	2, 8, 8				

The letters do not represent actual symbols of the elements. Complete the table by filling in the table

a) Atomic numbers of the elements (5 marks)
b) Electronic configuration of the atoms (5 marks)
c) Group and period of the elements (10 marks)
d) Atomic radius of P is greater than its ionic radius. Explain (2 marks)

e) Ionic radius of Y is greater than its atomic radius. Explain (2 marks)

f) Element U is more reactive than element V. Explain (2 marks)

g) Write formula of the compound formed between R and S

29. An isotope of element E has 34 neutrons and its mass number is 64. E forms a cation with 28 electrons. Write the formula of the cation with 28 electrons. Write the formula of the cation indicating the mass and atomic numbers.

30. Write electronic arrangement of the elements in brackets in the following compounds
    I. PCl_5 (P)

    II. PCl_3 (P)

    III. Cl_2O (Cl)

    IV. **Cl_2O_7 (Cl)**

CHAPTER THREE: STRUCTURE AND BONDING

Specific Objectives

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

a) describe the role of the outer electrons in determining chemical bonding
b) explain qualitatively the formation of covalent and ionic bonds
c) illustrate the covalent and ionic bonds using diagrams
d) explain the unique nature of the metallic bonding
e) state the effect of intermolecular forces of attraction on physical properties of substanccs
f) distinguish between bond types on the basis of physical properties of substances
g) compare and explain the changes in bond type across a period
h) select appropriate materials for use based on bond type.

INTRODUCTION

Definition of structure and bonding

Structure – This is the arrangement of particles.

Bonding – These are the forces holding the particles together.

TYPES BONDS AND RELATED STRUCTURE

There are three major types of bonds:

	BOND		**RELATED STRUCTURE**
1	Ionic bonding		Giant ionic
2	Metallic bonding		Giant metallic structure
3	Covalent	**Atoms**	Giant covalent/atomic structure
		Molecules	Simple molecular structure

Other bonds

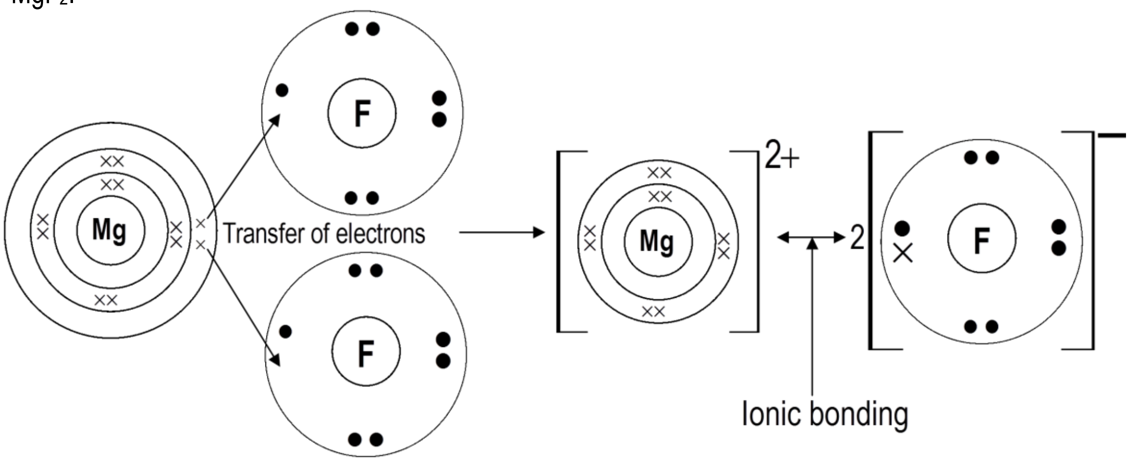
Dative /co-ordinate bond

Common mistake is that most students make when answering questions on ionic bonding is that they say a bond is ionic because it is formed between metal and non-metals; this **assumption is wrong** as some substances like $AlCl_3$ have covalent structures. The expected response in such question is ionic bonding involve **complete transfer of electrons.**

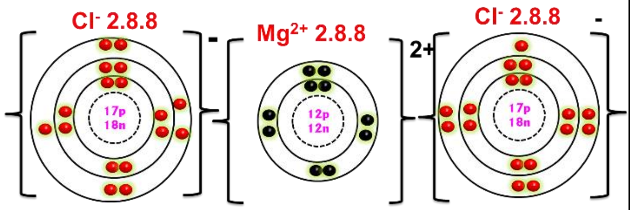
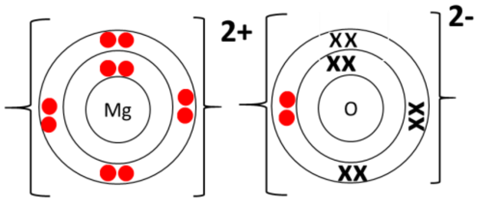
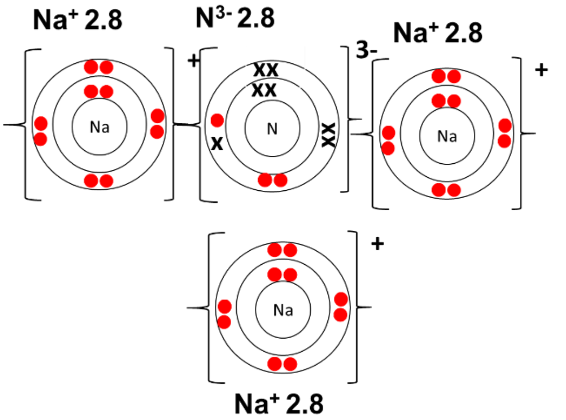
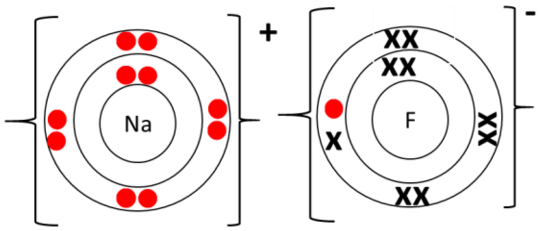
IONIC BONDING AND GIANT IONIC STRUCTURE

Points to note about ionic bonding

- Usually, Ionic bond is formed between metal and a non-metal.
- It involves ***complete transfer of electrons*** whereby the metal loses and the non-metal gains.
- The metal loses electrons to form a positively charged ion (cation) while the non-metal gains to form a negatively charged ion (anion).
- The resulting oppositely charged ions attract each other forming a strong ionic bonding E.g. formation of MgF_2.

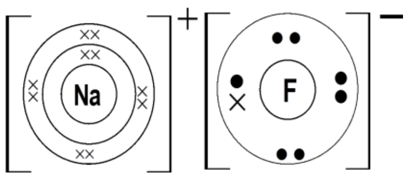
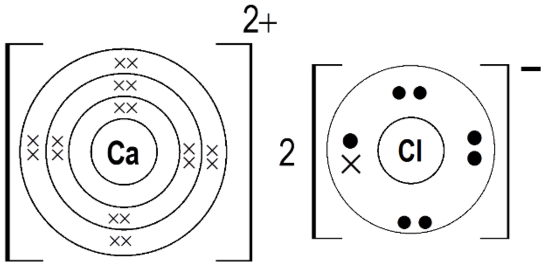
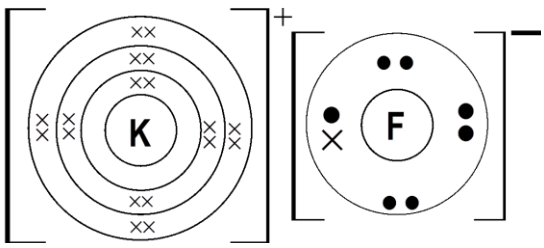


F

<table>
<tr><th>formation of Magnesium chloride</th><th>formation of Magnesium oxide</th></tr>
<tr><td>
</td><td>
</td></tr>
<tr><th>formation of sodium Nitride</th><th>formation of sodium Flouride</th></tr>
<tr><td>
</td><td>
</td></tr>
</table>

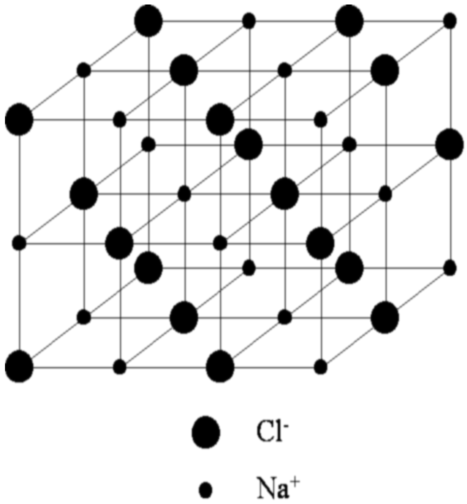
Points to note about ionic bond

- *All the electrons must be shown not just the outermost electrons*
- *Ionic bonding involves equal sharing of bonding electrons*
- *The best way to draw dot and cross diagram of ionic compound is make sure the like charges are not placed in the same point , a positive charged ion should be surrounded by negatively charged ions and negatively charged ion should be surrounded by positively charged ions.*
- *Avoid grouping of ions except in unavoidable cases like drawing of* Mg_3N_2 *and Aluminium oxide*

Compound	Dot and cross diagram	Compound	Dot and cross diagram
NaF		$CaCl_2$	
Na_2O		Al_2O_3	
MgO		KF	
CaS		Na_3P	
Mg_3N_2		Li_2S	

Na_2S		**MgF_2**	
NaOH		**$Mg(OH)_2$**	

GIANT IONIC STRUCTURES



Ionic structures are three dimensional in shape, e.g. in NaCl each Na^+ is surrounded by six Cl^- ions and each Cl^- is surrounded by six Na^+ ions which are equidistant and this leads to formation of a three dimensional structure.

Properties of giant ionic structures

They are usually solids at room temperature.

- They have high Melting point and boiling point due to strong ionic bonding.
- They are soluble in water but insoluble in organic solvents. They are made up of oppositely charged ions that are attracted by polar water molecules.
- They are crystalline.
- They are good conductors of heat and electricity in aqueous and molten state but not in solid state. This is because in aqueous or molten state the ions are mobile but in solid state the ion are fixed.

NB---*Common mistakes* *with students when asked to explain why ionic structure conduct electricity in aqueous but not in solid state, they state that* ***in solid state there are*** *no ions* ***but in aqueous state there are mobile ions/ions.***

"*No ions*" is wrong as the ions are there but are immobile/ fixed.

COVALENT BOND

Covalent bond is formed when there is sharing of bonding electrons.

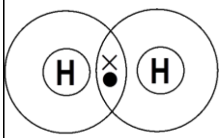
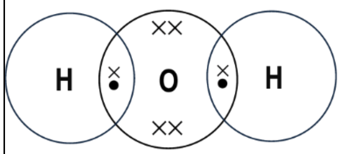
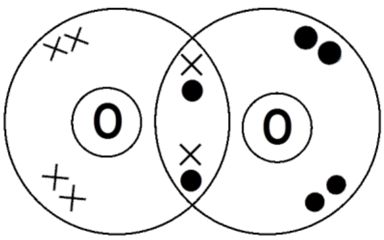
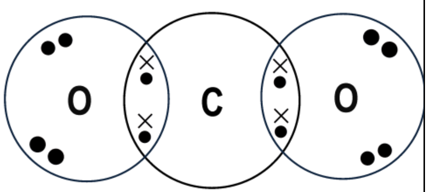
It is formed mainly when non-metals combine.

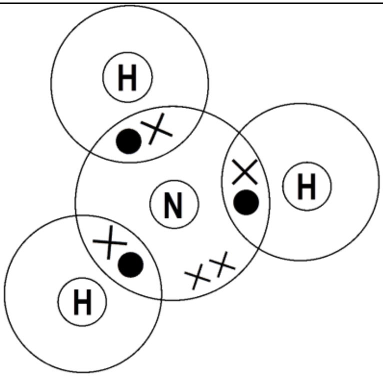
There are two types of covalent structures.

- ✓ Molecular structures
- ✓ Giant covalent/atomic structure

Molecular structures

Most substances with molecular structures are gases or liquids with low Melting point and boiling point .
In molecular structures, atoms share electrons such that each atom becomes stable e.g.

Compound	Structure	Compound	Structure
H_2		H_2O	
Cl_2		$SiCl_4$	
O_2		HCl	
CO_2		PCl_3	

CH_4		NH_3	
C_2H_4		C_2H_6	
N_2		PH_3	
H_2S		F_2	
CCl_4		Cl_2O	

NOCl(Nitrosyl Chloride)	Chloramine (NH_2Cl)

MOLECULAR STRUCTURES AND INTERMOLECULAR FORCES OF ATTRACTION

Molecular substances like hydrogen gas, chlorine, water, hydrocarbons, carbon (IV) oxide have simple molecular structures. The molecules are held together by intermolecular forces of attraction.
There are two intermolecular forces;

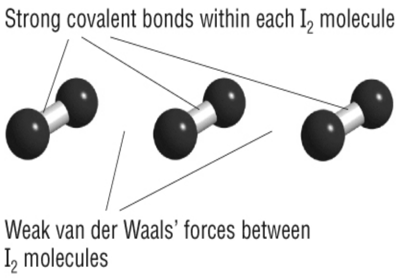
- ✓ Van der Waals forces
- ✓ Hydrogen bonding

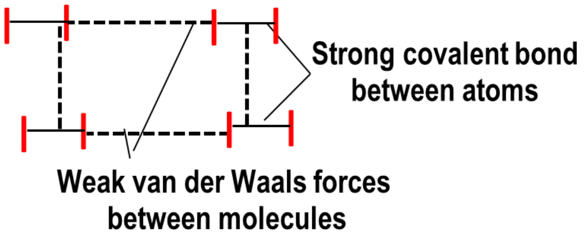
Some molecules are held together van der Waals forces only while some other molecules are held together by a combination of both van der Waals forces and hydrogen bonding.

Van der Waals forces

Molecular structures have two types of bonds; the **covalent bonds between the atoms** which are very strong and the **weak Van der Waals forces between the molecules** which are easily broken. That is why substances with weak van der Waals forces have low melting points and boiling points. The strength/number of van der Waals forces increase with increase in size of molecules and that's why melting points and boiling points of halogens increase down the group with increase in molecular mass/size.

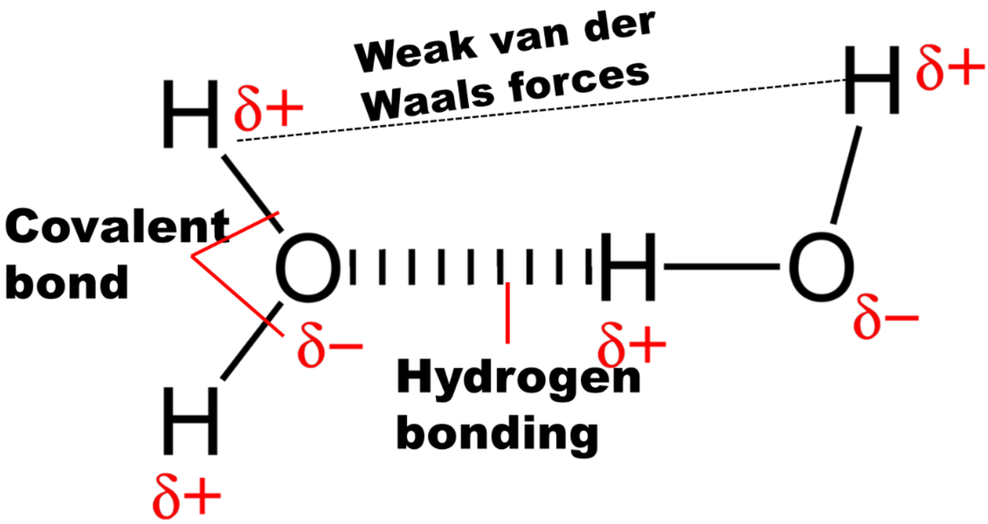
Structure of iodine molecule





Hydrogen bonding

This type of bonding is found in compounds containing ***hydrogen and a more electronegative element***. Hydrogen bonding is only form between hydrogen and Flourine and Oxygen .This results in the hydrogen atom being shared between two adjacent molecules. The hydrogen bonding is ***stronger than van der Waals forces and requires an extra energy to break*** hence compounds with hydrogen bonding have relatively higher melting and boiling point than their corresponding compounds with only van der Waals forces.
Van der Waals forces are weak forces of attraction between molecules or atoms which exist only when the particles are close to one another. Hydrogen bonding is stronger than the van der Waals forces but weaker than the covalent bond. Compounds with hydrogen bonding include water, ethanol sugar, methanol, ethanoic acid and sulphuric acid. **Compounds with hydrogen bonding are soluble in water.**



Points to note

*The strength **of van der Waals forces increases with increase** in* **molecular mass /size of molecule or atom***. That is why melting point and boiling point increase down the group in halogens and noble gases.*

It is also the reason why **phosphorous, sulphur and iodine** *exist as solids at room temperature despite having molecular structure, this is because they have relative large molecule hence have stronger van de Waals forces.*

Molecular structures with hydrogen bonding are soluble in water.

Properties of molecular structures

- Soluble in organic solvents but insoluble in water.
- Low melting point and boiling point due to weak van der Waals forces.
- Usually gases or liquids at room temperature due to weak van der Waals forces.
- Poor conductors of heat and electricity in both solid and molten state.
- Some molecular substances may, however, react with water to form solutions containing ions which do allow the flow of an electric current. E.g. HCl.

GIANT COVALENT/ GIANT ATOMIC STRUCTURE

This occurs in group IV elements and some of their compounds. In this type of structure, we have atoms which are covalently bonded to other atoms infinitely throughout the structure. Consequently, this type of structure is very strong and substances with type of structure are very hard solids which are usually used as abrasives.

Examples of giant covalent structures;

- Diamond
- Graphite
- Silicon
- Silicon (IV) oxide

Diamond and graphite are the two allotropes of carbon.

Allotropes are **different crystalline forms** of the same element in the same physical state.

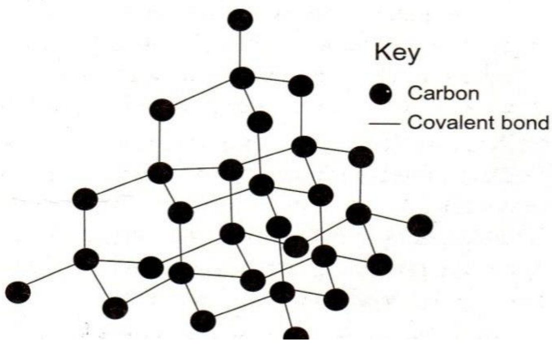
Allotropy is **the existence of different crystalline forms** of the same element in the same physical state.

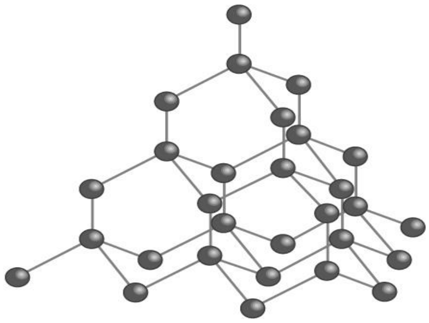
When defining an allotrope **remember** that the term **crystalline** is key to that definition. This is because it distinguishes the allotropes from other non-crystalline .e.g. amorphous carbons are not allotropes of carbon.

All elements which are allotropic have different melting points and boiling points, so if you are given an element with two melting points then that element exhibit allotropy.

DIAMOND

In diamond, each carbon atom is covalently bonded to four other carbon atoms forming a giant covalent structure. **In diamond all the valence electrons are involved in bonding leaving no free electron and that is why diamond does not conduct electricity.**





Properties of diamond

- It is colourless, transparent and shiny crystalline solid.
- It has a high density of 3.51g/cm^3 due to close packing of carbon atoms.
- It is the hardest substance due to uniformity of the strong covalent bonds throughout the structure.
- It has a high melting point and boiling point due to strong covalent.

Uses of diamond in relation to its structure

- For drilling and cutting other metals due to its hardness.
- For making ornaments/jewelry because it shines when polished.

GRAPHITE

Graphite is the other allotrope of carbon where each **carbon atom is covalently bonded to three** other carbon atoms leaving **one electron free** hence graphite conducts heat and electricity due to presence of **delocalized electrons.**

Graphite is made of hexagonal layers which can easily slide over each other and that is why graphite is used a lubricant

The hexagonal layers are held together by weak van der Waals forces and that is why graphite is soft with relatively lower boiling point than diamond.

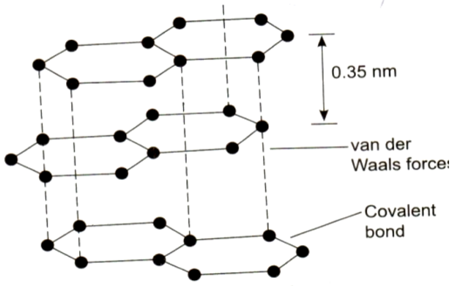
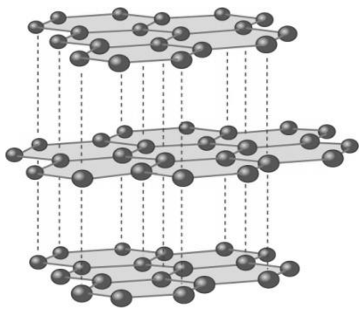


Fig 3.8: Structure of graphite



Reasons why graphite is used an electrode in electrolysis.

- Cheap
- Good conductor
- Inert

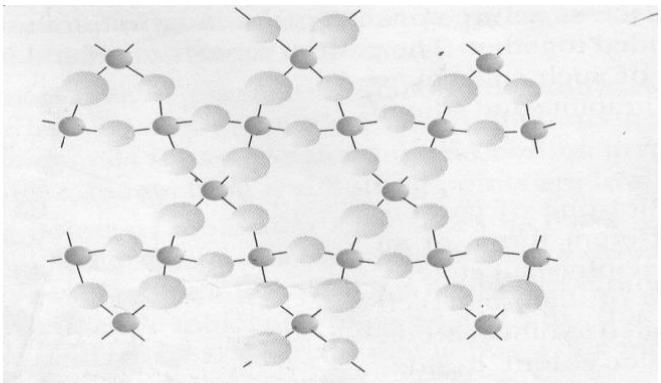
Uses of graphite in relation to its structure

- As an electrode in electrolysis because it conducts electricity.
- Make lead pencils because of the hexagonal layers which are held together by weak van der Waals forces that can easily peel off other **when pressed**.
- As a lubricant due to hexagonal layers held together by weak van der Waals which can easily slide over each other.

Differences between graphite and diamond

GRAPHITE	DIAMOND
Opaque	Transparent
Soft-slippery	Hard
Conductor	Non-conductor
Has both covalent and van der Waals	Has uniform covalent bonds
Has hexagonal layers	Has tetrahedral shape
Has relatively lower density	Has a higher density
Has relatively lower melting point and boiling point	Has relatively higher melting point and boiling point

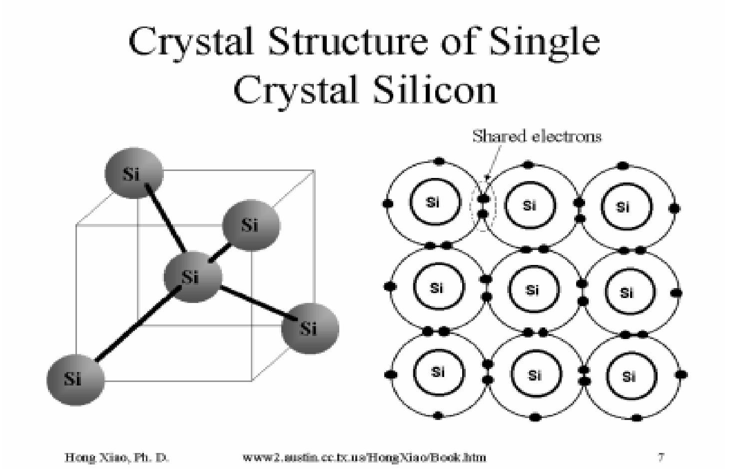
SILICON (IV) OXIDE

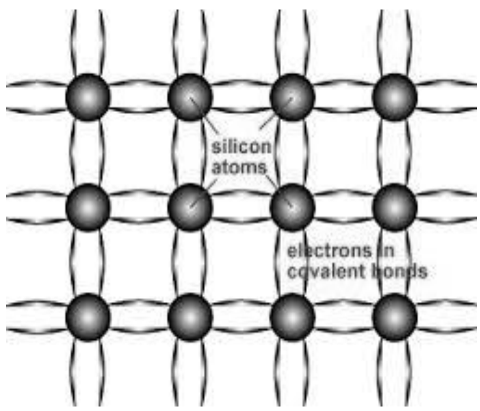


In silicon (IV) oxide, each silicon atom is covalently bonded to **four oxygen** atoms and each oxygen atom is covalently bonded to two silicon atoms. This results to a giant covalent structure with very high melting point and boiling point.

SILICON

Silicon in group IV period three and has a giant atomic structure, whereby each silicon is covalently bonded to other four silicon atoms forming a giant covalent structure with very strong covalent bond.





Properties of giant covalent structures

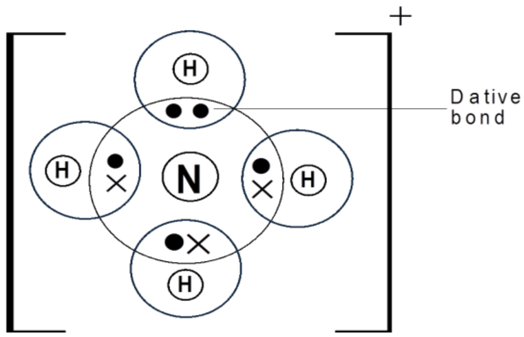
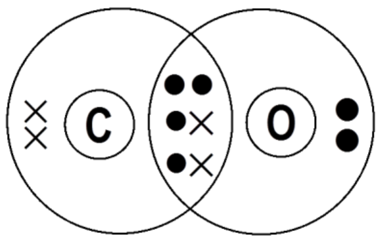
- They have high melting point and boiling point due to strong covalent bond.
- They are poor conductors of heat and electricity except graphite, silicon is a semi-conductor.
- They are insoluble in water.

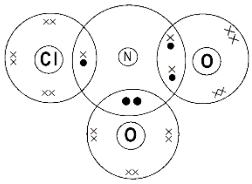
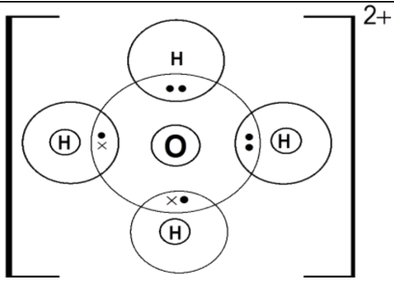
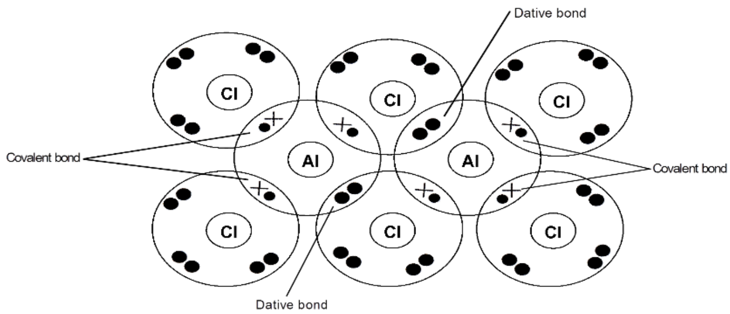
DATIVE/CO-ORDINATE BOND

This is a special type of covalent bond that involves sharing of a lone pair of electrons.

A lone pair means an extra pair of electron is donated by one atom only (unequal sharing).

E.g.

Compound	Structure
NH_4^+	
PH_4^+	
CO	
O_3	
H_3O^+	

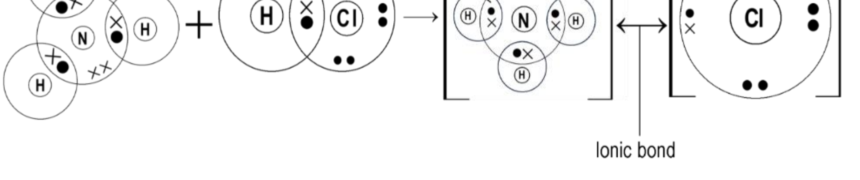
NO_2Cl	
H_4O^{2+}	
Al_2Cl_6	

Note:

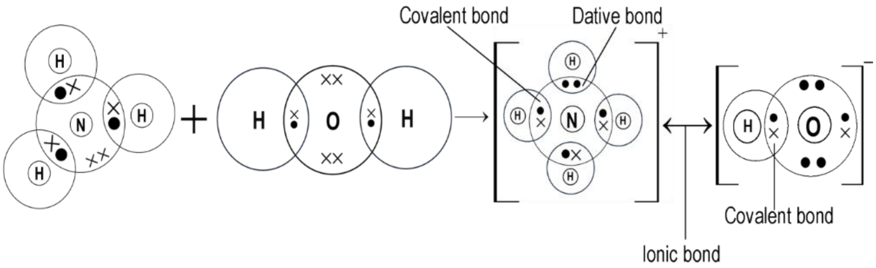
Some compounds have complex structures which comprise of ionic, covalent and dative bond in the same compounds. E.g. water molecule has covalent bond between O and H atom and hydrogen bonding whereby the hydrogen atom is bonded by two oxygen atoms between adjacent molecules. Water also has van der Waals forces between the molecules.

Compounds like ammonium chloride and ammonium hydroxide have covalent, ionic and dative bonds.

(a) $NH_3 \quad + \quad HCl \longrightarrow NH_4Cl$



(b) NH_3 + H_2O ⟶ NH_4OH

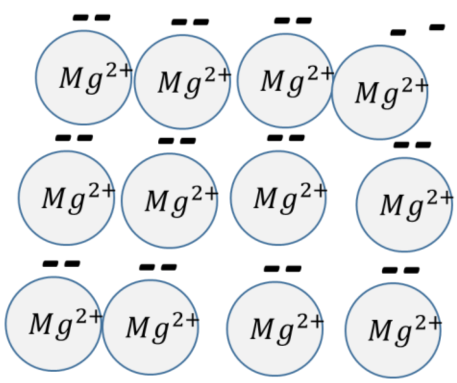


METALLIC BONDING

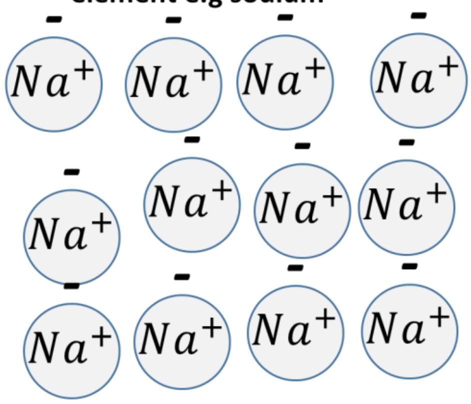
In a metallic structure the atoms are arranged like potatoes in a debe. The delocalized electrons are able to move freely in the metal ;lattice, the nuclei appear to be immersed in a sea of mobile electrons, this constitutes a mutual attraction between the delocalized electrons and the positive nuclei. This force of attraction is called electrostatic forces and leads to a strong metallic bond. The sea of electrons explains its electrical conductivity and its thermal conductivity.

The sea of electrons bonds the positive nucleus tightly into the lattice and this explains its high melting point. Since strong forces of attraction exists even in the liquid phase, metals tend to have a wide temperature range over which they remain liquid.

Metallic structure of group 2 element e.g Magnesium



Metallic structure of group 1 element e.g sodium



Properties of giant metallic structures

- high boiling point and melting point due to strong metallic bond

Note **the strength of metallic bond increases** with increase **<u>in number of delocalized electrons</u>** *and decrease in atomic radius. That is why melting point and boiling point decrease down the group in alkali metals and increase across the period three from sodium to Alluminium*

Electrical conductivity for metal whether in molten or solid states is due to presence of delocalized electrons

- good conductors of heat and electricity in both solid and molten state due to presence of delocalized electrons
- insoluble in water
- they are shiny
- they are malleable and ductile

Explain why the electrical conductivity of metals decreases with increase in temperature.

As the temperature increases the kinetic energy of the electrons increases and they move in a rapid random motion and the positive centers vibrate and this interferes with electrical conductivity,

Explain why the MP and BP increase from Na to Al

This is because across the period there is decrease in atomic radius and increase in number of delocalised electrons from Na to Al and this increases the strength of metallic bonding.

Comparison on structure and bonding.

Bond	Structure	M.P & B.P	Electrical conductivity		Solubility in water
			Aqueous /molten	Solid	
Metallic	Giant metallic	High	Good	Good	Insoluble
Ionic	Giant ionic structure	high	Good	poor	Soluble
Covalent	Giant atomic	high	Poor	poor	Insoluble
	Simple molecular	low	Poor	poor	Insoluble

Exceptions

1. *Mercury is metal but has a low melting point and boiling point and is a liquid at room temperature*
2. *Hydrogen chloride gas has a simple molecular structure but it is soluble in water and is a good conductor at aqueous state*
3. *Graphite has a giant atomic structure and is a good conductor of electricity at both solid and liquid state*
4. *Aluminium chloride has a simple molecular structure and not ionic and it is fairly soluble in organic solvents as well as in water*
5. *Silicon is a non metal but has a giant covalent structure with strong covalent bond and that makes it to have a very high melting point and boiling point*
6. *Given the Melting point and boiling point you should be able to determine which element is either a gas, solid or liquid at room temperature.(given particular temperature)*
7. *Iodine, sulphur and phosphorous have simple molecular structure but are solids at room temperature*
8. *Sugar ,water ,alkanols and alkanoic acids have hydrogen bonding in addition to van der waals forces and that makes them to have a higher melting point and boiling poing than the other molecular structures.*
    - *Substance is liquid at room temperature if the Melting point is below room temp and the boiling point is above room temperature(25^{o}C)*
    - *Substance is a solid at room temperature if the Melting point is above room temperature*
    - *Substance is a gas at room temperature if the boiling point is below the room temperature*
    - *Compare electric conductivity of metals and ionic compounds e.g. metal- delocalized electrons, ionic compounds –mobile ions.*
9. *For all the questions of chemistry whereby you are told to compare melting point and boiling point you have to invoke the concept of structure and bonding*
10. *For ionic compounds they only conduct electricity in <u>molten or aqueous state</u> due to presence of mobile ions .for metal and graphite electrical conductivity occur both in solid and liquid/molten state due to presence of delocalized electrons*

' ***POINTS TO NOTE WHEN ANSWERING QUESTIONS ON STRUCTURE AND BONDING***

1) *when told to compare the melting and boiling points or explain why a substance is as solid and another is a liquid or gas at a particular temperature you must invoke the concept of structure and bonding when answering the questions*
2) *strength of metallic bond is determined by the atomic radius and number of delocalized electrons ,the smaller the atomic radius the stronger the metallic bond that is why metallic bond decrease down the group and increase across the period in metals,*
3) *the melting point and Boiling point in molecular structures increase with increase in molecular mass due to increase in strength of van der waals forces*
4) *when comparing alkanols and alkanoic acids you say alkanoic acids have high Melting point and boiling point because they have stronger hvdrogen bonding than alkanols.*

examples

a). *Explain why CO_2 is gas at room temperature while SiO_2 is a solid at room temperature,*

CO_2 has a simple molecular structure with weak van der waals forces while SiO_2 has giant covalent structure with strong covalent bond.

b) *Compare the melting point of magnesium and sodium metal*

Magnesium has higher melting point than sodium, this is because magnesium has a stronger metallic bonding than sodium as Mg has a smaller atomic radius and more valence electrons than sodium.

c)*Explain why $MgCl_2$ is a solid at room temperature while $SICl_4$ is a gas at room temperature.*

$MgCl_2$ has a giant ionic structure with strong ionic bond while $SICl_4$ has simple molecular structure with weak van der waals forces

d)*Explain why iodine is a solid at room temperature while Chlorine is a gas at room temperature*

Iodine has stronger van der waals forces than chlorine due to its larger molecular size.

Nb. Strength of van der waals forces increase with increase in molecular size

Nb; Strength of ionic bond increases with increase in electronegativity and electropositivity of the combining elements e.g NaF has stronger ionic bond than NaCl because fluorine is more electronegative than chlorine

REVISION QUESTIONS ON STRUCTURE AND BONDING

1. **Name the type of bonds and related structure that exist in the following substances**
   Summary of bonds and related structure, complete the table below, more than one type of bond can exist in a /substance.

1)	Substance	Bonding (s)	Structure
2)	Hydrogen	**Covalent**	**Simple molecular**
3)	Helium		
4)	Lithium	**Metallic**	**Giant metallic**
5)	Carbon	**Covalent**	**Giant covalent**
6)	Nitrogen		
7)	Oxygen		
8)	Fluorine		
9)	Neon		
10	Sodium		
11	Magnesium		
12	Aluminium		
13	Silicon		
14	Phosphorous		
15	Sulphur		
16	Chlorine		
17	Argon		
18	Potassium		
19	Calcium		
20	Hydrogen chloride	Covalent	Molecular
21	Sulphur (IV) oxide	Covalent, dative	Molecular
22	Methane		
23	Water	Covalent,hydrogen bonding	
24	Sugar		
25	Methanol		
26	Methanoic acid		
27	Magnesium chloride	Ionic	Giant ionic
28	Sodium oxide		
29	Silicon (IV) chloride		
30	Magsium sulphide		
31	Ammonium ion		
32	Hydroxonium ion		

33	Aluminium chloride(dimer)		
34	**Ozone**		
35	**Sulphur (VI) oxide**		
36	**Carbon (IV) oxide**		
37	**Carbon (II) oxide**		
38	**Diamond**		
39	**Graphite**		
40	Phosphorous (III) chloride		

2. Explain why the boiling point of ethanol is 78°C while that of dimethyl ether is -24°C, although they have the same molecular mass. (2 marks).

3. Explain why water is a liquid at room temperature while methane is a gas at room temperature (2 marks).

4. Explain why melting and boiling point of halogens increase down the group while it decreases down alkali metals (2 marks)

5. Explain why chlorine and fluorine are gases, bromine is a liquid and iodine is a solid at room temperature (2 marks)

6. Explain why electrical conductivity of metals decrease with increase in temperature. (2 marks)

7. Explain why Aluminium has a high melting and boiling point than sodium (2 marks)

8. Ex plain why sodium is soft and magnesium is hard(2mk)

9. name the particles responsible for electrical conductivity in:
   I. Molten magnesium chloride(1mk)
   II. Molten magnesium(1mk)
   III. Magnesium metal(1mk)

10. Strength of metallic bond increases with decrease in atomic radius (2 marks)

11. The table below gives some information about elements represented by letters B, C, D and E. Study the information and answer the questions that follow:

Element	Atomic radii (nm)	Melting Point ^0C
B	0.152	180
C	0.186	98
D	0.231	64
F	0.244	39

i.Would these elements form part of group or period? Explain. (2 marks)

ii.What does the trend in melting points suggest about the nature of the elements (1 mark)

12. An element X has atomic number 30 while element Y has atomic number 8

a) Write the electron arrangement for X (1 mark)

b) What type of bond would be formed when X and Y react (1 mark)

13. Oxygen and Sulphur belong to group (V) of the periodic table. Explain why there is a big difference in their melting points (melting point of oxygen is – 216^0C while that of Sulphur is 44^0C.

14. The table below gives the distance between atoms (bond lengths) in halogen molecules and the energies required to break the bonds (bond energies) between the atoms.

Molecule	Bond length (nm)	Bond Energy (KJ Mol^{-1})
Cl – Cl	0.20	240
Br – Br	0.23	195
I – I	0.26	150
At – At	0.29	

a) Predict the energy required to break the At – At bond

b) What is the relationship between bond length and bond energy for the halogen molecules?

15. Using dots (.) and crosses (x) to represent outermost elections, draw a diagram to show the bonding in CCl_4 , PH_3 and Sulphur (IV) oxide

16. In terms of structure and bonding, explain why graphite conducts electricity while diamond does not (2 marks)

17. Using dots (.) and crosses (x) to represent electrons draw diagram to represent `the bonding in:

a) CO (1 mark)

b) $MgCl_2$ (1 mark)

18. State why an ammonia molecule (NH_3) can combine with H+ to form NH_4^+ (Atomic numbers: N=7 and H=1) (1 mark)

19. Draw the structure of phosphorous trichloride (PCl_3), PH_4^+ and H_4O^{2+} (3 marks)

20. Study the table below and answer the questions that follow:

Element	Atomic number	Relative atomic mass	Melting point (oC)
Aluminium	13	27.0	660
Calcium	20	40.0	842
Carbon		12.0	3550
Hydrogen		1.0	-259
Magnesium	12	24.3	650
Neon	10		-249
Phosphorous	15	31.0	44.2 (white),590 (red)
Sodium		23	98

a) Complete the table by filling in this missing atomic numbers and atomic mass. (2 marks)

b) Write the electron arrangement for the following ions (2 marks)

Ca^{+}

P^{3+}

c) What is the melting point of hydrogen in degrees Kelvin? (1 mark)

d) Which of the allotropes of phosphorous has a higher density? Explain (2 marks)

e) The mass numbers of three isotopes of magnesium are 24, 25 and 26. The relative atomic mass of magnesium is 24.243. What is the mass number of the most abundant isotope of magnesium? Explain (2 marks)

f) Give the formula of the compound formed between aluminium and carbon. (1 marks)

g) Explain the difference in the melting points of magnesium and sodium. (2 marks)

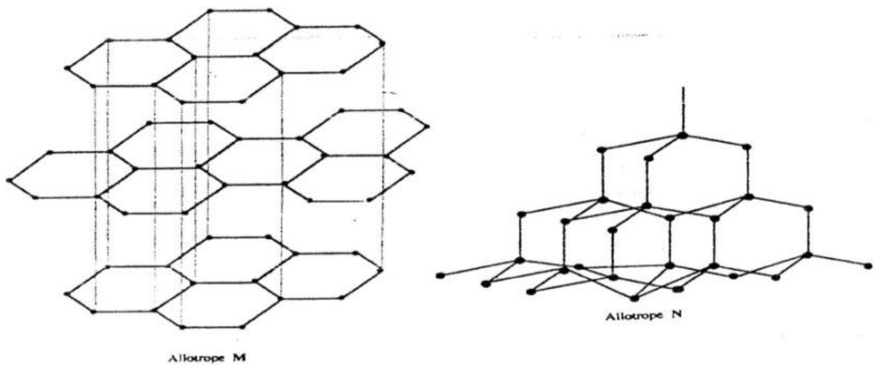
21. Diamond and graphite are allotropes of carbon. In terms of structure and bonding, explain the following.

a) Diamond is used in drilling through hard rocks

b) Graphite is used as a lubricant

22. Compound Q is a solid with a giant ionic structure. In what form would the compound conduct an electric current

23. The following diagrams show the structures of two allotropes of carbon. Study them and answer the questions that follow



a) Name allotrope (2 marks)

M ....

N...

b) Give one use of N (1 mark)

c) Which allotrope conducts electricity? Explain (2 marks)

24. With reference to iodine, distinguish between covalent bonds and Van Der Waals forces

25. Both diamond and graphite have giant atomic structures. Explain why diamond is hard while graphite is soft (3 marks)

26. Using dots (.) and crosses(x) to represent electrons, show bonding in the compounds formed when the following elements react: (Si = 14, Na = 11 and Cl = 17) (1 mark)

a) Sodium and chlorine (1 mark)

b) Silicon and chlorine (1 mark)

27. The table below gives some properties of substances I, II, III, and IV. Study it and answer the questions that follow

Substance	Electrical conductivity		M.P (^0C)	B.P (^0C)
	Solid	Molten		
I	Does not conduct	Conducts	801	1420
II	Conducts	Conducts	650	1107
III	Does not conduct	Does not conduct	1700	2200
IV	Does not conduct	Does not conduct	113	440

a) What type of bonding exists in substances I and II (2 marks)

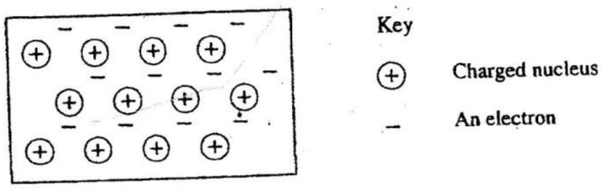
b) Which substance is likely to be sulphur? Explain (2 marks)

28. Distinguish between a covalent bond and a co-ordinate bond (2 marks)

29. In terms of structure and bonding, explain
    a) Why the boiling point of chlorine is lower than that of iodine. (2 marks)

    b) Why the silicon (IV) oxide is a solid while carbon (IV) oxide is a gas at room temperature (2 marks)

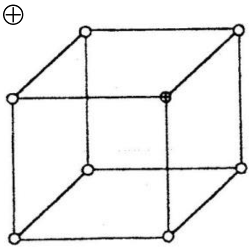
30. The diagram below is a section of a model of the structure of element T.



a) State the type of bonding that exists in T. (1 mark)

b) In which group of the period table does element T belong? Give a reason, (2 marks)

31. a). The diagram below represents part of the structure of a sodium chloride crystal. The position of one of the sodium ions in the crystal is shown as ⊕



i. On the diagram, mark the position of the other three sodium ions (2 marks)
ii. The melting and boiling points of sodium chloride are 801^0C and 1413^0C respectively. Explain why sodium chloride does not conduct electricity at 25^0C, but does so at temperatures between 801^0 C and 1413^0C (2 marks)

b). Give a reason why ammonia gas is highly soluble in water (2 marks)

32. The table below gives information about elements $A_1A_2A_3$, and A_4

Element	Atomic Number	Atomic Radius (nm)	Ionic radius (nm)
A_1	3	0.134	0.074
A_2	5	0.090	0.012
A_3	13	0.143	0.050
A_4	17	0.099	0.181

a) In which period of the periodic table is element A_2? Give a reason (2 marks)

b) Explain why the atomic radius of:

(i) A_1 is greater than that of A_2; (2 marks)

(ii) A_4 is smaller than its ionic radius (2 marks)

c) Select the element which is in the same group as A_3 (1 mark)

d) Using dots (.) and crosses(x) to represent outermost electrons. Draw a diagram to show the bonding in the compound formed when A_1 reacts with A_4 (1 mark)

33. In terms of structure and bonding, explain why the melting point of oxygen is much lower than that of sodium.(3 marks)

34. Using dots (.) and crosses (X) ,show bonding in:

a) The compound formed when nitrogen reacts with fluorine (Atomic numbers F= 9, N=7)

b) Sodium oxide (Atomic numbers Na = 11,O= 8)

35. a). Using electrons in the outermost energy level, draw the dot (**.**) and cross (x) diagram for the molecules H_2O and C_2H_4. (H = 1, C = 6, O = 8) (2 marks)

i. H_2O

ii. C_2H_4

b). The formula of a complex ion is $Zn(NH_3)_4^{2+}$. Name the type of bond that is likely to exist between zinc and ammonia in the complex ion.

36. 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 (^0C)**
R	11	97.8
S	12	650.0
T	15	44.0
U	17	-102
V	18	-189
W	19	64.0

a) Give the reasons why the melting point of:

i) S is higher than that of R (2 marks)

ii) V is lower than that of U

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) Give one use of element V (1 mark)

37. The table below gives the number of electrons, protons and neutrons in particles A, B, C, D, E, F and G.

Particle	Protons	Electrons	Neutrons
A	6	6	6
B	10	10	12
C	12	10	12
D	6	6	8
E	13	10	14
F	17	17	18
G	8	10	8

i. Write the formula of the compound formed when E combines with G. (1 mark)

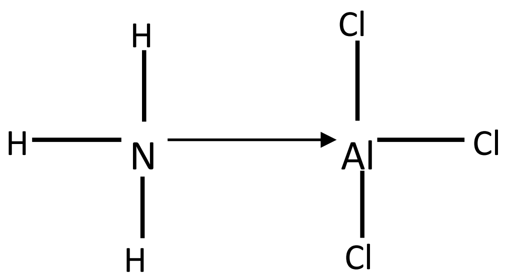
ii. Name the type of bond formed in (iii) above (1 mark)

iii. How does the radii of C and E compare? Give a reason. (2 marks)

iv. Draw a dot (.) and cross (x) diagram for the compound formed between A and F. (1 mark)

v. Why would particle B not react with particle D? (1 mark)

38. The diagram below shows the bonding between aluminum chloride and ammonia.

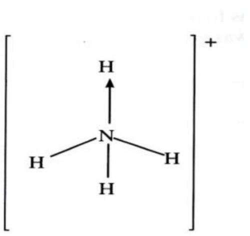


a) Name the types of bonds that exist in the molecule (1 mark)

b) How many electrons are used for bonding in the molecule? (1 mark)

c)

39. Ammonium ion has the following structure:



Label on the structure

a) Covalent bond; (1 mark
b) Coordinate (dative) bond (1 mark)

CHAPTER FOUR: TRENDS IN PROPERTIES ACROSS PERIOD THREE

The elements in period 3 are Sodium, Magnesium, Aluminium, Silicon, Phosphorous, Chlorine, Argon

They each have three occupied energy levels and hence belong to period three.

The elements show a gradual change in both their physical and chemical properties

A) TRENDS IN PHYSICAL PROPERTIES

Table showing summary of physical properties of period three elements

Element	Proton number	Atomic radius	Appearance	Bond	structure	Electrical conductivity	Melting point (K)	Boiling point(K)
Sodium	11	0.190	Grey solid	Metallic	Giant metallic	Good	371	1156
magnesium	12	0.160	Grey solid	Metallic	Giant metallic	Good	922	1380
Aluminium	13	0.125	Grey solid	Metallic	Giant metallic	Good	933	2740
Silicon	14	0.117	Black solid	covalent	Giant atomic	Semi-condctor	1683	2628
phosphorus	15	0.110	White solid Red solid	covalent	molecular		44 (white) 590(red)	280
Sulphur	16	0.104	Yellow solid	covalent	molecular	Poor	113 119	718
Chlorine	17	0.099	Green-yellow gas	covalent	molecular	Poor	172	238
Argon	18	0.197	Colourless gas	covalent	molecular	Poor	84	87

Trends in physical properties across period three

i)The atomic radius

The atomic radius **decreases** across the period due to **increase in number of protons** that leads to increase in nuclear force of attraction of the outermost electrons towards the nucleus hence decreasing the atomic radius, therefore sodium has the largest atomic radius.

ii)Trend in bond type

Sodium, Magnesium and Aluminium have **giant metallic structure with** strong metallic bonding.

- Silicon has giant atomic/covalent structure with strong covalent bonding
- Phosphorous, Sulphur, chlorine and argon have simple molecular structure with weak van der waals between molecules.

iii)Melting Points and Boiling Points

Sodium, magnesium and Aluminium

Sodium, magnesium and Aluminium are all metals. They have metallic bonding with giant metallic structure, in which positive metal ions are attracted to delocalized electrons. Going from sodium to Aluminium:

,the atomic radius decreases and the number of delocalized electrons increases from 1 to 3 ... so the strength of the metallic bonding increases across the period hence melting points and boiling points increase.

Silicon

Silicon is a metalloid (an element with some of the properties of metals and some of the properties of non-metals). Silicon has a giant covalent structure with strong covalent bonding. in which each silicon atom is covalently-bonded to four other silicon atoms therefore Silicon has a very high melting point and boiling point because:

All the silicon atoms are held together by strong covalent bonds.

Phosphorus, Sulphur, Chlorine and Argon

These are all **non-metals**, and they exist as small, separate molecules. Phosphorus, Sulphur and chlorine exist as **simple molecules**, with strong covalent bonds between their atoms. Argon exists as separate atoms (it is **monatomic**).

Their melting and boiling points are very low because:

- The molecules are held together by van der Waals' forces between the molecules which are very weak bonds and easily broken so little energy is needed to overcome them and they have low melting points and boiling points.

Sulphur has a higher melting point and boiling point than the other three because:

- Phosphorus exists as **tetratomic** P_4 molecules, Sulphur exists as **octaatomic** S_8 molecules , chlorine exists as **diatomic** Cl_2 molecules and argon exists as **monatomic** Ar atoms . the **strength/number** of the **van der Waals' forces** decreases as the molecular mass/ size of the molecule decreases and therefore the **melting points and boiling points decrease** in the order $S_8 > P_4 > Cl_2 >$ **Ar**

element	atomicity of the molecule	molecuar mass
Sulphur	Octa atomic (s_8)	256
Phosphorus	Tetra atomic (p_4)	128
chlorine	diatomic (cl_2)	71
Argon	Monoatomic	40

Sample practice questions

1. **Explain why chlorine has a higher melting point than argon**
   - Chlorine is **diatomic** molecule while argon is **monatomic**; therefore chlorine has **stronger/more** van der waals forces than argon
2. **Explain why phosphorous is a solid while chlorine is a gas at room temperature**
   - Phosphorous is **tetra atomic** while chlorine is **diatomic** ,hence phosphorous having a **higher** molecular mass has **stronger/more** van der waals forces than chlorine.
3. **Explain why sulphur has a higher boiling point than phosphorous**
   - Sulphur exist as octatomic molecule while phosphorous exist as tetratomic , therefore sulphur has a **stronger/more** van der waals forces than phosphorous
4. **Explain why magnesium has a higher melting point than sodium**
   - Magnesium has a **stronger metallic bonding** than sodium because it has **more** delocalized electrons than sodium
5. **compare the reactivity of magnesium and sodium with chlorine**
   - Sodium reacts more vigorously with chlorine than magnesium, this is because sodium has a larger atomic radius hence more electropositive /loses its outer electron ***more*** easily than magnesium.
6. **Elements P,Q.S have atomic numbers 11,12,and 13 respectively .which element has the highest melting point, explain**
   - This is because **S** has the **smallest atomic radius** and **highest number of delocalized** electrons hence the **stronges**t metallic bonding.
7. **Explain why silicon has the highest melting point and boiling point**
   - It has a giant covalent structure with strong covalent bonding in which each silicon atom is covalently-bonded to four other silicon atoms, which requires a lot of energy to break

Common errors with students

- When answering questions on bonding or reactivity, or comparing melting point and boiling points, the following must be considered.
- If comparison involves two elements then use a comparative(second degree of comparison) term like more easily, stronger, lesser ,more reactive ,higher melting point, more delocalized e.t.c refer to quiz 1-5 above , if those comparative terms are not used the candidate will not score. For example is a candidate writes magnesium has a strong metallic bonding than sodium ,he/she will not score ,the correct term is stronger not strong
- If comparison involves more than two elements of substances use a superlative e.g highest strongest/ greatest/least/smallest ,if such terms are not used then the candidates does not score refer to quiz 6
- When comparing metallic bonding it is wrong to say metallic bonding increase with increase in valence electrons ,the correct term should be more delocalized electrons not valence electrons this is because not all elements with valence electrons form metallic bonding

iv)Electrical Conductivity Across the Period

Sodium, magnesium and Aluminium are good conductors of electricity due to presence of delocalized electrons but conductivity increase across the period due to increase in number of delocalized electrons

- Silicon is a semi-conductor
- Phosphorus, Sulphur, chlorine and argon are poor conductor as they don't have delocalized electrons.

B:TRENDS IN CHEMICAL PROPERTIES

Metals

Sodium, Magnesium and Aluminium are metals and their reactivity decrease across the period due to decrease in atomic radius and increase in nuclear charge which increase the attraction between the outermost electrons and nucleus. The number of electrons increase hence more electrons require more energy to remove, hence sodium having the largest atomic radius and loses one electron to become stable is the most reactive while Aluminium has the least atomic radius and loses three electrons to become stable is the least reactive

SILICON

Silicon has four electrons in the outermost energy levels.

Usually silicon does not form ionic compounds because it will require a lot of energy to lose or gain four electrons, therefore forms compounds through covalent bonding,

PHOSPHORUS, SULPHUR AND CHLORINE,

Chlorine is the most reactive, followed by phosphorous then sulphur, this is attributed to decrease in forces holding the molecules together. Therefore chlorine having relatively smaller atomic radius and requiring only one electron to be stable is the most reactive.

Argon has a stable configuration and therefore does not react

i) Reaction with oxygen

Sodium and Magnesium; refer to chemical families and air and combustion

Aluminium

Aluminium when heated in oxygen react to form Aluminium oxide is unreactive, this oxide layer prevents any further reaction, that's why Aluminium is used to make cooking vessels like sufurias

$$4Al_{(s)} + 3O_{2\,(g)} \longrightarrow 2Al_2O_{3(s)}$$

Non metals

Silicon, Phosphorus and sulphur react with oxygen to form corresponding oxides but chlorine and argon don't react with oxygen.

Phosphorus burns in air with a white flame forming a mixture of Phosphorous (V)oxide and Phosphorous (III)oxide, the oxides dissolve in water forming acidic solutions.

$$Phosphorus \quad + \quad Oxygen \longrightarrow Phosphorous\ (V)oxide$$

$$P_{4(s)} + 5O_{2(g)} \longrightarrow P_4O_{10(s)}$$

Phosphorus + Oxygen ⟶ Phosphorous (III)oxide

$$P_{4(s)} + 3O_{2(g)} \longrightarrow 2P_2O_{3(s)}$$

Sulphur burns in oxygen with a **blue flame** forming *Sulphur (IV)oxide*

$$Sulphur \quad + \quad Oxygen \longrightarrow Sulphur(IV)oxide$$

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

Silicon + Oxygen ⟶ Silicon (IV)oxide

$$Si_{(s)} + O_{2(g)} \longrightarrow SiO_{2(g)}$$

ii) Reaction with Water

Sodium and magnesium refer to chemical families

Aluminium does not react with cold water or steam due to presence a coating of aluminium oxide, which prevents any further reaction.

Non-metals do not react with water therefore Silicon, Phosphorus and sulphur don't react with water. But chlorine reacts with water forming chlorine water which is a mixture of hydrochloric (HCl) acid and choric (I) acid (Hypochlorous acid). (HOCl)

$$Cl_{2(g)} + H_2O_{(l)} \longrightarrow HCl_{(aq)} + HOCl_{(aq)}$$

iii) Reaction with acids

Magnesium reacts with both dilute hydrochloric acid and dilute sulphuric (VI) acid forming salt and hydrogen gas

$$Mg_{(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_{2(g)}$$

$$Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$$

Aluminium does not readily react with dilute acids **due to presence a coating of Aluminium oxide, however on removing the coating of Aluminium the metal reacts**

$$2Al_{(s)} + 3H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 3H_{2(g)}$$

$$2Al_{(s)} + 3HCl_{(aq)} \rightarrow AlCl_{3(aq)} + 3H_{2(g)}$$

The reaction between sodium and dilute acids is explosive and should never be tried. Silicon, Phosphorus, silicon and sulphur do not react with dilute acids

TRENDS IN STRUCTURE AND BONDING OF COMPOUNDS OF PERIOD THREE ELEMENTS

Summary in properties of oxides

Oxides	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_4O_6, P_4O_{10}**	**SO_2**
Structure	Giant ionic structure			Giant atomic	Simple molecualr	
Bonding	Ionic bond			Covalent	Weak van-der waals forces of attraction	
M.P	1280	2900	2040	1610	24, 580	-17
Acidic/basic (nature)	Basic		amphoteric	acidic		
Reaction with water	Alkaline solution pH =12-14	Weakly alkaline pH 9-10	Insoluble pH 7		Acidic pH 1 or 2	

i) Bonding and trends in melting point of the oxides

1. Oxides of Na, Mg and Al form **<u>giant ionic</u>** structures with **strong ionic bonding**
   - 👉 Large amount of energy is needed to overcome the **<u>strong electrostatic forces</u>** of attraction **between the ions** and thus they have high melting points.
2. SiO_2 forms **<u>giant covalent</u>** structure
   - 👉 Large amount of energy is needed to overcome the **<u>strong covalent bonding</u> between the Si and O atoms** and thus SiO_2 has high melting point.
3. Oxides of P and S
   - 👉 These form **<u>simple molecular</u>** structures with molecules held together by **<u>weak Van der Waals' forces</u>**, small amount of energy is needed to overcome the **<u>weak Van der Waals' forces</u>** of attraction **between the molecules** and thus they have low melting points.

ii)Nature of oxides across period three

- Na_2O. MgO are basic oxides
- Al_2O_3 is amphoteric and react with both acids and alkalis
- SiO_2 , P_2O_5, and SO_2 are acidic and react with alkalis

iii)Solubility in water

- Na_2O and MgO dissolve in water to form alkaline solutions; Na_2O is very soluble while MgO is slightly soluble in water.
- Al_2O_3 does not dissolve in water
- Oxides of non-metals dissolve in water to form acidic solution; however, silicon (IV) oxide does not dissolve in water

$$Phosphorous\ (V)\ oxide + Water \longrightarrow\ Phosphoric\ (V)\ acid$$

$$P_2O_{5(s)} \quad + \ 3H2O_{(l)} \ \longrightarrow 2H_3PO_{4(aq)}$$

$$Phosphorous\ (III)oxide + Water \longrightarrow Phosphoric(III)acid$$

$$P_2O_{3(s)} \quad + \ 3H2O_{(l)} \ \longrightarrow \ 2H_3PO_{3(aq)}$$

$$Sulphur\ (IV)\ oxide + \quad Water \ \longrightarrow \ Sulphuric\ (IV)\ acid$$

$$SO_{2(g)} \ + \ H_2O_{(l)} \longrightarrow \ H_2SO_{3(aq)}$$

$$SiO_{2(g)} \ + \ H_2O_{(l)} \longrightarrow H_2SiO_{3(aq)}$$

iv)Reaction with acids

- Na_2O and MgO being basic react with acids ,to form salt and water
- Al_2O_3 react with both acids and alkali therefore is said to be amphoteric oxide.
- Oxides of non-metals do not react with acids but react with alkalis

CHLORIDES OF ELEMENTS IN PERIOD THREE

Trends in properties in bond types and properties Chlorides of elements of period 3

Chlorides	**NaCl**	**$MgCl_2$**	**ACl_3**	**$SiCl_4$**	**PCl_5**	**SCl_2**
Physical state	Solid	Solid	solid	liquid	liquid	gas
M.P (^oC)	801	714	sublimes	-70	sublimes	-78
B.P (^oC)	1437	1412	sublimes	57	sublimes	59
Structure	Giant ionic	Giant ionic	molecular	molecular	molecular	molecular
Bond type	Ionic	Ionic	covalent	covalent	covalent	covalent
pH of solution	7	7	2	3	2	3
Solubility on water	Soluble	Soluble	hydrolyses	hydrolyses	hydrolyses	hydrolyses

Solubility in water and nature of solutions

- NaCl, $MgCl_2$ have giant ionic structure and dissolve in water to form neutral solutions
- $AlCl_3$ have simple molecular structure and is soluble in water and fairly soluble in organic solvents
- $AlCl_3$ undergo hydrolysis in water to form acidic solution,
- **NB**. Salt of ions with charge density of +3 and above e.g. Al^{3+}, Fe^{3+} **undergo hydrolysis in water to form acidic solution. That is why when Na_2CO_3 is added to a solution of Al_2SO_4/$AlCl_3$ / $FeCl_3$ there is effervescence and also when a blue litmus paper is inserted it turns from blue to red**.

PCl_3 PCl_5 , $SiCl_4$ dissolve in water forming acidic solutions

$$PCl_{3(s)} + 3H_2O_{(l)} \longrightarrow 3HCl_{(g)} + H_3PO_3$$

$$PCl_{5(s)} + 4H_2O_{(l)} \longrightarrow 5HCl_{(g)} + H_3PO_{4(aq)}$$

$$SiCl_{4(l)} + 2H_2O_{(l)} \longrightarrow SiO_{2(s)} + 4HCl_{(g)}$$

The chlorides of phosphorous and silicon fume in air as they hydrolyse in presence of moisture to form HCl gas which appear as white fumes

Trends in bond types and melting point and boiling points of chlorides

NaCl and $MgCl_2$ have giant ionic structure with strong ionic bond hence have high melting point and boiling point. NaCl have higher boiling point than $MgCl_2$ because sodium has a larger atomic radius and therefore is more electropositive than magnesium hence forms a stronger ionic bond than magnesium.

$AlCl_3$ forms a dimer molecule whereby the molecules are held together by weak van der waals hence easily broken require less energy to break, hence low melting points and boiling points.

$SiCl_4$, PCl_5 and SCl_2 have molecular structures with weak van der waals forces hence have low melting point and boiling points.

SUMMARY OF TRENDS IN PROPERTIES OF ELEMENTS DOWN AND ACROSS PERIODIC TABLE

1) Down the group

a. Atomic radius

- For all elements the atomic radii increase down the group due to increase in number of occupied energy levels down the group

b. Atomic mass

- For all elements the atomic mass increase down the group due to increase in number of protons and neutrons down the group

c. Reactivity

Trend in reactivity down the group depends on whether it is a metal or non-metal

I. For metals

Reactivity: Increases down the group due to the weakening of attraction between the positive nuclei and the outer most electrons with increase in atomic radii.

Ionization energy

Decease down the group due to decrease in attraction between the positive nucleus and outermost electrons with increase in atomic radius.

Electropositivity

Increase down the group due to decrease in attraction between the positive nucleus and outermost electrons with increase in atomic radius

Note: strongest reducing agent/ most electropositive element/most reactive metal /element with least ionization energy -this is an element that easily loses electron. Here you select the metal with the largest atomic ***radius and with fewer electrons in the outermost energy level. Mostly group one***

II. **For non metals**

Reactivity of non metals decreases down the group due to decrease in effective nuclear attraction and shielding effect with increase in atomic radii/ number of occupied energy levels.

d. **Melting point and boiling point:**

I. **For metals**

M.P and B. P decreases down the group due to decrease in strength of metallic bonding with increase in atomic radii, increase in atomic radii leads to decrease in attraction between outermost electrons and positive nucleus.

II. **For non metals**

Melting point and boiling point increases down the group due to increase in strength of Vander Waals forces with increase in molecular mass/atomic radius.

Electron affinity

Electron affinity decrease down the group with increase in atomic radius hence increased shielding effect .

2. ACROSS THE PERIOD

a. **Atomic radius**

Generally for all elements the atomic radius decreases across the period due to increase in number of protons that attracts the outermost electrons more strongly towards the nucleus, hence decreasing the atomic radius.

a. **Reactivity**

In Metals

Reactivity decreases across the period due to increase in the number of electrons to be lost and also decrease in atomic radii.

Ionization energy

Increase across the period *due to increase in attraction between the positive nucleus and outermost electrons with decrease in atomic radius and increase in nuclear charge.*

Electropositivity

Decrease across the period *due to increase in attraction between the positive nucleus and outermost electrons with decrease in atomic radius and increase in nuclear charge.*

Melting and boiling point increase across the periods due to increase in the strength of metallic bond with increase in delocalized electrons and decrease in atomic radius.for non metals it depends on the size of the molecule eg. S›P›Cl›Ar

Electric conductivity increase from sodium to Aluminium due increase in number of delocalized electrons. Silicon is a metalloid while the rest of non metals are poor conductors of electricity.

d) ***Strongest oxidizing agent/ most electronegative element /most reactive non-metal*** /**-,** this is usually the non-metal with the smallest atomic radius and requires fewer electrons to be stable, mostly halogens

Comparison between structures

- All metals have giant metallic structures with strong metallic bond
- All compounds of metals have giant ionic structure with strong ionic bond except Aluminium chloride which has simple molecular structure with weak van der waals forces
- All non-metals and their compounds have simple molecular structure with weak van der waals forces except silicon and silicon (IV) oxide which have giant covalent structure with strong covalent bond

Examination content

- Compare giant metallic structure vs simple molecular structure
- Compare giant covalent structure vs simple molecular structure
- Compare giant ionic structure vs simple molecular structure
- Compare simple molecular structures with hydrogen bonding and those without
- Comparing simple molecular structures vs molecular structure (look size of molecule ,atomicity of molecule
- Comparing metallic bond (look for atomic radius (for same group or number of delocalized electrons (for same period)

REVISON QUESTIONS ON MIXED CONCEPTS

This area tests on structure of atom, structure and bonding chemical families and trends across the period

1. Study the table below to answer the questions that follow. The letters do not show the actual symbols of the elements.

Element	B	C	D	E	F	G	H	I	J	K
Atomic number	7	8	19	15	2	9	6	16	12	11
Atomic mass	14	16	39	31	4	19	12	32	24	23
Melting point	-	-	63.7	44	-272	-223	vary	113	669	98

a) Select two elements with oxidation states of -2 (1 mark)

b) Which elements represent

i. the most powerful oxidizing agent (1 mark)

ii. the most powerful reducing agent(1 mark)

c) Which element has the highest ionization energy (1 mark)

d) Select two elements when reacted form a compound that conduct electricity in both molten and aqueous state.(1 mark)

e) Select two elements when reacted form a compound that dissolves in water to form acidic solution (1mk)

f) Using dots (.) and crosses (X) to represent electrons draw diagrams showing bonding between B and J (1mk)

g) Explain why for some elements the mass number is not twice the atomic number (1 mark)

h) Explain why the melting point of element K is higher that of element D (2 marks)

i) Describe how the solid mixture of sulphate of D and $PbSO_4$ can be separated (3 marks)

element	Atomic number	Melting point
L	11	97.8
K	12	650
M	13	660
N	14	1410
Q	17	-101
R	19	63.7

2. Study the information in the table below and answer the question that follow. the letters do not represent the symbols of the elements.

a) Write the electronic arrangement for ions formed by elements M and Q
M……………………………………
Q……………………………………….
b) Select an element which is:
   i. the most electronegative element
   ii. a poor conductor of electricity
c) In which period of the periodic table does element Q belong? Explain
d) Compare the re-activity of element R and L. Explain your answer

e) Using dots (•) and crosses (x) to represent outmost electrons; show bonding in the compound formed by elements M and Q.

f) Explain why the melting point of element M is higher than that of element L

g) Write an equation for the reaction that would occur between K and water

h) In terms of structure and bonding explain why there is a large difference in the melting points of N and other elements.

3. Study the table below representing part of periodic table and answer the questions that follow. The letters do not represent the actual symbols of the elements

F			*P*		*H*	*G*	*H*	I
	Q		*J*	*K*		*L*	*M*	
N		*X-Z*						

a) What type of bond would you expect in the compound formed between H and F, explain (2 marks)

b) Which of the elements J and M will have a greater atomic radius? Explain.(2 marks)

c) Elements F and N are in the same group of periodic table, how do their atomic radius compare? Explain (2 marks)

d) An element W has atomic number 15. indicate the position it would occupy in the table above (2 marks)

e) What name is given to elements X-Z (2mks)

f) Why is J used in electric cables where Q is not (2 marks)

g) Element K is termed as a metalloid, what does the term metalloid mean (2 marks)

h) How would you expect reactivity of H and M to compare (2 marks

i) Compare the atomic mass of N and F (2 marks)

4. (a) The table below shows the atomic numbers of elements of the periodic table represented by letter J to Q. The letters are not the actual chemical symbols for the elements

Element	J	K	L	M	N	P	O	Q
Atomic number	2	7	8	10	11	12	13	14

(i) Select two elements which belong to:
I The same period of the periodic table (1 mark)

II The same group of the periodic table

(ii) Select the element which (1 mark)
I Will form a divalent anion

II Reacts most vigorously with water

(iii) Has the greatest tendency to form covalent bond. Explain(2mks)

(b). The boiling points of some chloride are shown in the table below:

Group	I	II	III	IV	V	VI	VII
Chloride	LiCl	Chloride of W	BCl3	CCl4	NCl3	OCl2	FCl
B.P. (0C)	1350	487	12	77	71	2	-101
Chloride	Nacl	Mgcl2	AlCl3	Chloride of X	PCL3	SCl2	Cl2
B.P. (0C)	1465	1418	Sublimes at 180	57	74	59	-35

(i) What is the most likely formulae for the chlorides of W and X?

(ii) Select two chlorides from the table which are the most ionic. Explain why the two selected chloride are the most ionic.

(iii) Would you expect group VIII elements of the periodic table to form chlorides? Explain the answer (1 mark)

5. Metal p is a group 2 element in the periodic table and it lies below Q in the same group

a) Explain how the reactivity of metal P and Q with bromine compares (1 mark)

b) Given that the atomic number of Q is 12, determine the atomic number of P. Show how you arrive at your answer (2 marks)

6. Element E has atomic number 15.

(a) Write the electron arrangement for an atom (1 mark)

(b) Explain why E forms a chloride which is a liquid of low boiling point (1 mark)

7. Element J whose atomic numbers is 31 has two isotopes. The table below shows the mass numbers and the relative abundance for each isotope

Mass number	Relative abundance (%)
69	60.4
71	39.6

(a) Determine the number of neutrons in the isotope with mass 69 (1 mark)

(b) Calculate the relative atomic mass of element J (2 marks)

Element	Atomic No.	M.P.0(C)	B.P. 0(C)	Ionic radius (nm)
A	11	98	890	0.095
B	12	650	1110	0.065
C	13	660	2470	0.050

D	14	1410	2360	0.41
E	15	44.2, 590	280	0.034, 0.212
F	16	113, 119	445	0.184
G	17	-101	-35	0.181
H	18	-189	-186	

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

a) i). Write electronic arrangement for the atoms represented by letters B and F (2 marks)

ii). State the nature of the oxides of the elements represented by B and F (1 mark)

b) Why does the element represented by letter E have two values of melting points? (1 mark)

c) Explain the following observations in terms of structure and bonding

i. There is an increase in boiling point from A to C

ii. Element D has a high boiling point

iii. F has a higher boiling point than G.

d) Explain the difference in ionic radius between elements represented by letters A and G (2 marks)

e) Write the formulae and the electronic arrangement of the two ions E whose ion radius are shown in the table

9. Complete the table below

Isotope	**Number of**		
	Protons	Neutrons	Electrons
$^{59}_{27}$**K**			

10. The electron arrangement of ions X^{3+} and Y^{2-} are 2, 8 and 2, 8, 8 respectively

(a) Write the electron arrangement of the elements X and Y (2 marks)

(b) Write the formula of the compound that would be formed between X and Y **(1 mark)**

11. With reference to its atomic number of one, explain why hydrogen can be placed in either group I or VII of the period table (1 mark)

12. The table below shows some properties of substances E, F, G and H

Substance	Action with water	Melting point	Thermal conductivity
E	Un reactive	High	Poor
F	Reactive	High	Poor
G	Unreactive	High	Good
H	Unreactive	Low	Good

Select the substance that would be most suitable

a) For making a cooking pot (1 mark)

b) As a thermal insulator (1 mark)

13. 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 number	9	10	11	12

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

b) Which two elements would react most vigorously with each other? (1 mark)

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

14. An element Y has the electronic configuration 2.8.5

a) Which period of the periodic table does the element belong? (1 mark)

b) Write the formula of the most suitable anion formed when element Y ionizes (1 mark)

c) Explain the difference between the atomic radius of element Y and its ionic radius (1 mark)

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

Element	Atomic number	Melting point	Formula of chloride	Melting point of chloride
G	11	98	GCl	801
H	12	650	HCl_2	715
J	14	1410	JCl-80_4	-70
K	16	113	K_4Cl_2	-80
L	20	851	LCl_2	780

a) Which elements are metals? Give a reason (2 marks)

b) Write the formula of the compound formed when element H reacts with element K (1 mark)

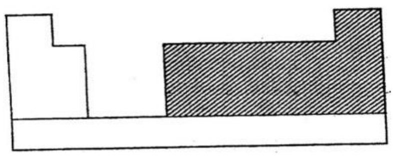
c) Explain why the melting point of J is higher than that of K (2 marks)

d) What is the oxidation state of J in its chloride? (1 mark)

e) How does the melting point of the fluoride of G compare with that of its chloride? Explain (2 marks)

f) Reactivity of H and L with water compare? Give an explanation (2 marks)

16. The chart below is an outline of part of the periodic table.



a) i). With the help of vertical and horizontal lines, indicate the direction of increasing metallic nature of the elements (2 marks)
ii). Which types of elements are represented in the shaded area? (1 mark)

b) Element A is the same group of the periodic table as chlorine. Write the formula of the compound formed when A reacts with potassium metal.

17. Study the information in the table and answer the questions that follow

Ion	Electronic arrangement	Ionic radius
Na^+	2.8	0.095
K^+	2.8.8	0.133
Mg^{2+}	2.8	0.065

a) Explain why the ionic radius of (1 mark)

i. K^+ is greater than that of Na^+

ii. Mg^{2+} is smaller than that of Na^+

18. An oxide of element F has the formula F_2O_5

a) Determine the oxidation state of F (1 mark)

b) In which group of the periodic- table is element F? (1 mark)

19. Explain why the reactivity of group (VII) elements decreases down the group

20. . The table below shows the first ionization energies of elements B and C.

Element	Ionisation energy KJ mol-1
B	494
C	736

What do these values suggest about the reactivity of B compared to that of C? Explain (2 marks)

21. Four metal F, G, H and J were each separately added to cold water, and steam. The table below is a summary of the observations made and the formulae of the hydroxides formed.

Metal	Cold water	Hot water	Steam	Formula of Hydroxide
F	Reacts slowly	Reacts fast	Reacts very fast	F(OH)2
G	No reaction	No reaction	No reaction	-
H	Fast	Reacts very fast	Reacts explosively	HOH
J	No reaction	Reacts slowly	Reacts fast	J(OH)2

a) Which two elements are likely to be in the same group of the periodic table? (2 marks)

b) Arrange the metals in the order of their reactivity starting with the most reactive (2 marks)

22. State two factors which determine the stability of an isotope (2 marks)

23. The table below shows properties of chlorine, bromine and iodine. Complete the table by giving the missing information in (i), (ii), and (iii) (3 marks)

Element	Formula	Colour and state room temperature	Solubility
Chlorine Bromine Iodine	Cl_2 Br_2 I_2	i)………………… Brown liquid iii) ……………….	Soluble ii)…………….. Slight soluble

24. 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	B	C	D	E	F
Atomic number	18	5	3	5	20
Mass number	40	10	7	11	40

a) Which two letters represent the same element? Give a reason (2 marks)

b) Give the number of neutrons in an atom of element D (1 mark)

25. The table below gives some information about elements I, II, III and IV which are in the same group of the periodic table. Use the information to answer the questions that follows.

Element	First Ionisation energy ($kjmol^{-1}$)	Atomic Radius (nm)
I	520	0.15
II	500	0.19
III	420	0.23
IV	400	0.25

State and explain the relationship between the variations in the first ionization energies and the atomic radii. (3 marks)

26. a). An atom Q can be represented as

$$^{52}_{24}Q$$

What does the number 52 represent? (1 mark) b). Study the information in the table below and answer the equations that follow. Letters are not actual symbols

Element	Electronic Arrangement of stable ion	AtomicRadius(nm)	IonicRadius(nm)
N	2.8.8	0.197	0.099
P	2.8.8	0.099	0.181
R	2.8	0.160	0.065
S	2.8	0.186	0.095
T	2	0.152	0.068
U	2.8	0.072	0.136

i. Write the formula of the compound formed when N reacts with P. (atomic numbers are N = 20; P = 17) (1 mark)

ii. Identify the elements which belong to the third period of the periodic table. Explain (2 marks)

iii. Which of the element identified in b (ii) above comes first in the third period? Explain (2 marks)

iv. Select two elements which are non- metals (1 mark)

27. a). The atomic numbers of elements C and D are 19 and 9 respectively. State and explain the electrical conductivity of the compound CD in:

i. Solid state (1 ½ marks)

ii. Aqueous state (1 ½ marks)

b). The table below gives the number of electrons, protons and neutrons in particles **A**,**B**,**C**,**D**,**E**,**F**, and **G**

Particle	Protons	Electrons	Neutrons
A	6	6	6
B	10	10	12
C	12	10	12
D	6	6	8
E	13	10	14
F	17	17	18
G	8	10	8

i. Which particle is likely to be a halogen? (1 mark)

ii. What is the mass number of E? (1 mark)

iii. Write the formula of the compound formed when E combines with G? (1 mark)

28. 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		**B**	**C**	
D			**E**	**F**			**G**	
							H	

a) Which is he 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) What is the name given to the group of elements where, **C**,**G** and **H** belong? (1 mark)

d) 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)

e) The melting points of elements **F** and **G** are 1410^0C 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)

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

g) **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)

29. The grid below shows a part of the periodic table. The letters do not represent the actual symbols. Use it to answer the questions that follow:-

C								
	K					U		T
X	Y		M			Q	W	
J								Z

a) i). How do the atomic radius of element **X** and **Y** compare ii). Using crosses (**X**) to represent electrons, draw the atomic structure of element **Q**

iii). State the period and the group to which element **Q** belong

b) i). The ionic configuration of element **G** is 2.8 **G** forms an ion of the type G^{-1}. Indicate on the grid, the position of element **G.**

ii). To which chemical family does element **G** belong?

c) i). State **one** use of element **U**

ii). What is the nature of the compound formed between **K** and **U?**

d) Select the element with the smallest atomic radius

30. Study the table below and answer the questions that follow.

Particle	Atomic number	Ionic configuration	Formula of oxide	Atomic radii	Ionic radii
P	4	…………..	……………	0.110	0.031
Q	…………..	2.8.8	QO	0.200	0.099
R	……………	2.8.8	R_2O	0.230	0.133
S	17	2.8.8	S_2O_7	0.099	0.181
T	16	……………	………	0.104	0.231

i. Complete the table above

ii. From the table, choose the most reactive metal. Explain

iii. Which element is the most electronegative. Explain

iv. Using dots (**.**) and crosses (x) to represent electrons, show the bonding in the chloride of **Q**

v. Explain the solubility of element **T** in water

vi. Why is Aluminium used to make utensils yet it is a reactive metal?

vii. Distinguish between valency and oxidation number

31. The grid below represents part of the periodic table. Study it and answer the questions that follow:

S				R	E		X	
Q	Z					M	T	V

a) Identify the element that gains electrons most readily

b) Which of the metal is most reactive? Explain

c) What name is given to the family of elements to which elements **X** and **T** belong?

d) Explain why:-

i. Ionic radius of **Q** is smaller than that of **M**

ii. Atomic radius of **Q** is greater than that of **S**

e) Which of the element in the table does not have the ability to form an ionic or covalent bond? Explain

f) Give the formula of the compound formed between **R** and **Z**

32. The grid below is part of the periodic table. The elements are not represented by their actual symbols. Use the information to answer the questions that follow.

T						K	S	
			W	R				N
Q								

a) Which is the most reactive:
   (i) Non — metal? Explain

   (ii) Metal? Explain

b) Name the family to which elements **T** and **Q** belongs.

c) Write the formula of the compound formed when **W** reacts with **S**.

d) Name the type of bond and structure formed when elements **R** and **K** react.

e) Explain why element **N** doesn't form compounds with other elements.

f) Compare the atomic radii of **T** and **Q.** Explain

33. Study the data given in the following table and answer the questions that follow. The letters are not the actual symbols of elements.

Element	Number of protons	Melting point	Bpt oC
A	11	98	890
B	12	650	1110
C	13	60	2470
D	14	1410	2360
E	15	442, 590	280
F	16	113, 119	445
G	17	-101	-35
H	18	-189	-186

a) State and explain the trend in melting point in **A B C**

b) Explain why the melting point and boiling points of element **D** are the highest

c) Explain why the element represented by letter **E** has two melting point values

d) Write down the chemical formula between element **C** and sulphate ions

e) Name the chemical family in which **H** belong and state one use of the element

f) What is the nature of the oxide of the elements represented by letters **C** and **F**?

34. The table below gives information on four elements by letters **K, L, M** and **N**. Study it and answer the questions that follow. The letters do not represent the actual symbol of the elements.

Element	Electron arrangement	Atomic radius (nm)	Ionic radius (nm)
K	2.8.2	0.136	0.065
L	2.8.7	0.099	0.181
M	2.8.8.1	0.203	0.133
N	2.8.8.2	0.174	0.099

a) Which **two** elements have similar properties? Explain

b) What is the most likely formula of the oxide of **L**?

c) Which element is non-metal? Explain

35. Study the information given below and answer the questions that follow:

Element	Atomic radius (nm)	Ionic radius (nm)	Formula of oxide	Melting point of oxide (ºC)
A	0.364	0.421	A2O	-119
D	0.830	0.711	DO2	837
E	0.592	0.485	E2O3	1466
G	0.381	0.446	G2O5	242
J	0.762	0.676	JO	1054

a) Write the formula of the compound formed when J combined with G

b) Explain why the melting point of the oxide of E is higher than that of the oxide of G

36. Study the information in the table below and answer the questions that follow:

Element	Atomic radius (nm)	Ionic radius (nm)
W	0.114	0.195
X	0.072	0.136
Y	0.133	0.216
Z	0.099	0.181

a) Would these form part of a metallic or a non-metallic group? Explain

b) Suggest an element in the table above likely to be the most reactive. Explain

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

Element	Electronic configuration	Boiling point
X	2.7	-188^oC
Y	2.8.7	-35^oC
Z	2.8.8.7	59^oC

a) What is the general name given to the group in which the elements **X,** Y and **Z** belong?

b) Select **two** elements which are coloured gases

c) Explain why **Z** has the highest boiling point

d) Write an equation for the reaction of element **Z** with iron metal

e) Element **Y** was dissolved in water and a piece of blue litmus paper was put into the resulting solution. State and explain the observation that was made on the litmus paper

38. The table below shows elements **A, B, C, E, F**, and **G**. Elements in group **X** have a valency of 2 while elements in group **Y** have a valency of 1. Use the table to answer the questions that follow:-

	GROUP X				**GROUP Y**	
Element	**A**	**B**	**C**	**E**	**F**	**G**
Atomic radius (nm)	14.0	19.5	19.7	5.2	7.9	11.3
Ionic radius (nm)	7.6	10.5	12.4	12.6	16.1	19.6

a) Atomic radius increases from **A** to **C** and from **E** to **G.** Explain.

b) Explain the difference in the atomic and ionic radii of group **X** elements

c) Elements **C** and **G** belong to the same period. Explain why the atomic radius of **C** is greater than that of **G**

d) Give the formula of the compound formed when **B** and **F** react

e) What type of bonding is formed in the compound above? Explain

f) Starting with the least reactive, arrange the elements in group **Y** in the order of reactivity. Explain:

39. The table below gives elements represented by letters T, U, V, W, X, Y and their atomic numbers

Element	T	U	V	W	X	Y
Atomic number	12	13	14	15	16	17
Electron arrangement						

Use the information in the table to answer the questions below: -

a) Complete the above table giving the electron arrangement of each of the elements

b) In which period of the periodic table do these elements belong? Give a reason

c) How does the atomic radius v compare with that of X? Explain

d) Give the formula of the compound that could be formed between U and W

e) What type of bonding will be present in a compound formed between T and Y? Explain

f) Arrange the species T^-, T^+ and T in increasing order of size

g) Which of the ions X^{2+} and X^{2-} is the most stable? Explain

h) Give the formula of:

I. An acidic oxide formed when one of the elements in the table is heated in air

II. A basic oxide formed when one of the elements in the table is heated in air

40. The grid below represents part of the periodic table. The letters do not represent the actual symbols of the 1. The grid below represents part of the periodic table. The letters do not represent the actual symbols of the elements. Study it and answer the questions that follow:

L							L	
M	P			T		J	U	X
N	Q		R	S		B	V	Y
H							W	
K								

\

a. Explain why element **L** appears in two different groups in the grid above (1mk)

b. i). State the name of the chemical family to which **P** and **Q** ,X and Y belong (1 mk)

the compound formed between **P** and **V** (1mk

iii). Name the bond type of compound above,give a reason

c.

i Compare the melting points of, N and K,S and B, V and W Explain(6mks)

ii Compare the atomic radius of H and W (2mks)

d.

i Select the element with the smallest and largest atomic radius ,explain (2mks

ii Identify an element whose oxide dissolves in both acids and alkalis (1 mark

iii Write the equation for the burning of **T** in excess air (1 mark iv Using dot and cross (**x**) to represent electrons, draw a diagram to illustrate bonding in sulphide of **Q** and chloride of R (2 marks)

v State **one** use of element **X** (1 mark)

vi Draw a well diagram show how gas J ,L and gas V can be prepared in the laboratory 6marks

vii Draw a well labelled diagram to show how you can prepare chloride of Q and Chloride of R by direct synthesis .

e. Excess powder of element Q was heated with 2.7875g of lead (II) oxide ,until there was no further change in mass . calculate the mass of metal formed and mass of metal R than reacted with lead (II) oxide

f. i). Select the most electronegative /strongest oxidizing /most reactive metal and the most electropositive/ strongest reducing element/most reactive metal iii). Compare the melting point and boiling point of the oxide T and S

iv). Give two commercial uses of element J (2 marks)

g. Select an element that:

i exist as Diatomic molecule

ii exist as Monoatomic gas

iii exist as Tetratomic molecule

iv exist as Octatonic molecule

v That has a simple molecular structure but is a solid at room temperature . give a reason for such

vi That forms a soluble carbonate

vii That exhibits allotropy

h. i). Compare the atomic mass of H and M (2 marks

ii). Compare the reactivity of H and M (2mks)

Compare the first ionization energy of H and M AND n AND Q(2mks) iii). Element Z is in the third period and forms ions with charge of -3, place it in the grid above

iv). Element Z forms two ions, write the formula and configuration of the ions

v). Elements P, Q,R ,and S have atomic numbers 2,14,18 and 20 respectively select two elements in the same chemical family

k. compare the rate of diffusion of gas v and W

.

i. Write the configuration of phosphorus in the following
    i. PH_3
    ii. PO_3^{3-}
    iii. PO_4^{3-}

j. Write the configuration of
    i. Ca^{2-}
    ii. Al^{+2}
    iii. O-

j. Compare the melting points of the following

i. Chloride of H and chloride of S

ii. Oxide of Q and oxide of B

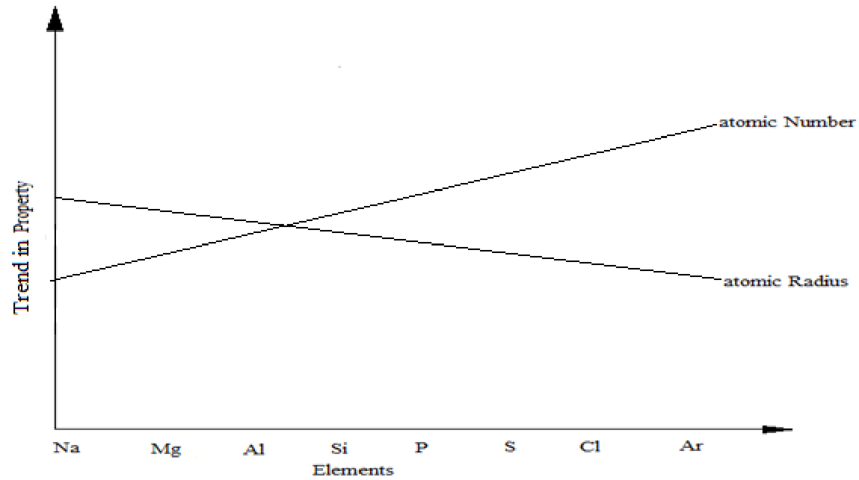
iii. Chloride of Q and chloride R

J.

i **Draw the atomic structure of element Q given that its mass number is 24 .**

ii **Write down the electronic configurations of the most stable ions of elements Q and** V

41. The figure below represents trends of some properties of period three elements. Study it answer the questions that follow.



k. Explain the trends shown by the atomic numbers and the atomic radii

(i) Atomic number (1 mark)

(ii) Atomic radii (2 marks)

42. 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.

G								
					I			V
K	L		M					
J								

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 ionization energies of two elements K and M at random are 577 kj/mol and 494kj/mol.

i).write equations for the 1st ionization energies for elements K and M and indicate their energies. (1 mark)

ii). Explain the answer in (b) (i). (2mark)

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 above . Explain (2 mark)

ii) How do the reactivity of elements J and K compare? Explain (2mark)

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.

Element	Formula of chloride	Nature of chloride solution
L		
M		

ii). The chloride of element M vaporizes easily while its oxide has a high melting point. Explain .

L. On the same axes, sketch the trend of reactivity across the period (1 mark)

Complete the table below

Compound	***nature of solution***
NaCl	
$MgCl_2$	

$AlCl_3$	
NaO	
P_2O_3	

Additional questions copy in your note books and answer

43. Study the table and answer the questions that follow.

Element	Atomic number	Atomic radius (nm)	Ionic radius (nm)	Melting point (^oC)
A	12	0.136	0.065	650
B	14	0.118	–	1410
C	17	0.099	0.181	–101
D	20	0.171	0.099	850

(a) Identify **two** elements with similar chemical properties . Explain. (2 marks)
(b) Explain why **A** has a large atomic radius than **B**. (1 mark)
(c) State the observations made when element **A** is placed in a beaker filled with water. (2 marks)
(d) Explain the difference in melting points of element **B** and **C**. (2 marks)
(e) How would you expect the pH values of aqueous solutions of the oxides of **C** and **D** to compare. Explain. (2 marks)
(f) Describe how crystals of sodium sulphate can be prepared starting with sodium metal. (3 marks)
(g) State **one** use of element **D**. (1 mark)

44. Study the information below and answer the questions that follow.

Formula of the chloride	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_3	SCl_2	
M.P (℃)	801	714	–	–70	–91	–80	
Formula of the oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_2	Cl_2O_7
M.P (^oC)	1190	3080	2050	1730	560	–73	–90

(a) Aluminium chloride $AlCl_3$, has an unexpected bond type and structure.

(i) State the type of bond and the structure in $AlCl_3$. (1 mark)
(ii) What type of bonding would $AlCl_3$ be expected to have why? (1 mark
(iii) Why is the melting point of $AlCl_3$ not indicated in the table above? (1 mark)

(b) A piece of blue litmus paper is placed in a solution of sodium chloride and a solution of aluminium chloride. Explain what would be observed in each case.

(i) Sodium chloride solution (1 mark)
(ii) Aluminium chloride solution (2 marks)

(c) Explain the large difference in the melting point of the compound of formula MgO and P_4O_{10}. (2 marks)

(d) Write down the equations for the reaction between the compounds of formula Na_2O and water. (1 mark)

(e) Silicon (IV) chloride gets hydrolyzed by water . Write a balanced equation for this reaction. (1 mark)

45. A natural element represented by letter Y has two types of atoms. The composition of the particles is as summarized below.

Type of atom	Nucleons present	% composition
$^{63}_{29}Y$	29, 34	____
$^{65}_{29}Y$	29, ____	30.9

(a) Complete the missing numbers. (1 mark)
(b) What is the name assigned to these two types of atoms? (1 mark)
(c) Which atom has the least percentage of abundance? (½ mark)
(d) Calculate the relative atomic mass of Y. (2 marks
(e) Explain what is meant by nuclear particles giving examples where possible. (1½ marks

46.

(a) Study the information given below and answer the questions that follow.

Element	Atomic radius (nm)	Ionic radius (nm)	Formula of oxide	Melting point of oxide (‘C)
P	0.364	0.421	A_2O	-119
Q	0.830	0.711	BO_2	837
R	0.592	0.485	E_2O_3	1466
S	0.381	0.446	G_2O_5	242
T	0.762	0.676	JO	1054

(i) Which elements are non-metals? Give a reason. (2 marks)
(ii) Explain why the melting point of the oxide of R is higher than that of the oxide of S. (2 marks)
(iii) Give **two** elements that would react vigorously with each other. Explain your answer. (2 marks)

(b) Study the information in the table below and answer the questions that follow (The letters do not represent the actual symbols of the elements)

		Ionization energy (kJ/mole)	
Element	Electronic configuration	1^{st} 1.E	2^{nd} 1.E
A	2.2	900	1800
B	2.8.2	736	1450
C	2.8.8.2	590	1150

(i) What chemical family do the elements A, B and C belong? (1 mark)
(ii) What is meant by the term ionization energy? (1 mark)
(iii) The 2^{nd} ionization energy is higher than the 1^{st} ionization energy of each. Explain. (1 mark)
(iv) When a piece of element C is placed in cold water, it sinks to the bottom and an effervescence of a colourless gas that bums explosively is produced. Use a simple diagram to illustrate how

this gas can be collected during this experiment.
(3 marks)

47. Study the part of the periodic table below and answer the questions that follow.
(The letters do not represent actual chemical symbols of the elements)

A			C		D			F
Q	B						E	

(a) Which element represents: (4 marks)
- i) Alkali metal
- ii) Halogen
- iii) Most reactive metal
- iv) Has an octet in the outermost energy level.

(b) Write the electronic arrangement of the: (4 marks)
- i) ion of B
- ii) ion of E
- iii) ion of A
- iv) atom of C

(c) An element X is in Group V and period 3. Indicate the position of X on the grid and write its electron arrangement in the same grid.
(2 marks)

(d) Explain why the atomic radius of D is smaller than that of C. (3 marks)

(e) Explain why the ionic radius of A is smaller than its atomic radius. (3 marks)

(f) Explain why the ionic radius of E is greater than its atomic radius. (3 marks)

48.
(a) The grid below represents part of the periodic table. Study it and answer the questions that follow.

				Q				
W						**R**	**S**	
T							**U**	
V							**Z**	

(i) Which element will require the least amount of energy to remove a valence electron? Explain. (2 marks)

(ii) Which one is the most reactive nonmetal? Explain. (2 marks

(iii) Write the formula of the compound formed between W and R. (1 mark)

(iv) Explain why the atomic radius of element S is smaller than that of R. (2 marks)

(b) Study the information in the table below and answer the questions that follow.

Formula of compounds	**NaCl**	**$MgCl_2$**	**$AlCl_3$**	**$SiCl_4$**	**PCl_3**	**S_2Cl_2**
Boiling point °C	**1470**	**1420**	**Sublimes at 180°C**	**60**	**75**	**60**
Melting point °C	**800**	**710**		**-70**	**-90**	**-80**

(i) Give two chlorides that are liquid at room temperature. (1 mark)

(ii) Which of the chloride would remain in liquid state for the highest temperature range (show how you arrive at your answer)
(2 marks)

(iii) Write chemical equations for the reaction of chloride of sulphur and phosphorus with water

(iv) . (4 marks)

(v) Draw a dot (●) and cross (**x**) diagram for aluminium chloride at 180 ºC. (2 marks)

49.

(a) The table below shows some of the properties of period III elements.

Element	Q	R	V	X	Y	Z
Atomic radii (nm)	0.136	0.099	0.125	0.117	0.110	0.157
Formula of oxide	QO	R_2O	V_2O_3	XO_2	Y_2O_5	Z_2O
Melting point (ºC)	650	110	660	119	44.2	97.8
Conductivity	Good	Poor	Good	Poor	Poor	Good

(i) From the table which elements would be:

I. Magnesium (1 mark)

II. Sulphur (1 mark)

(ii) Write the formula of the chloride of X. (1 mark)

(iii) Arrange the elements as they appear from left to right of periodic table. (2 marks)

(iv) A solution of aluminium chloride is acidic to litmus. Explain this observation. (1 mark)

(v) What type of bonding would you expect the chloride of Y to exhibit? (1 mark)

(vi) Write a chemical equation for the reaction of oxide of Y with water. (2 marks)

(b) Use the information in the table below to answer the questions that follow.

	Ionization energy in kJ			
Element	1st	2nd	3rd	4th
V	320	580	4900	7200
W	430	4300	6200	8300
X	600	1050	1450	16000
Y	7000	9000	10500	14000
Z	200	420	5200	6600

(i) Name two elements that can be found in the same group? Explain. (2 marks)

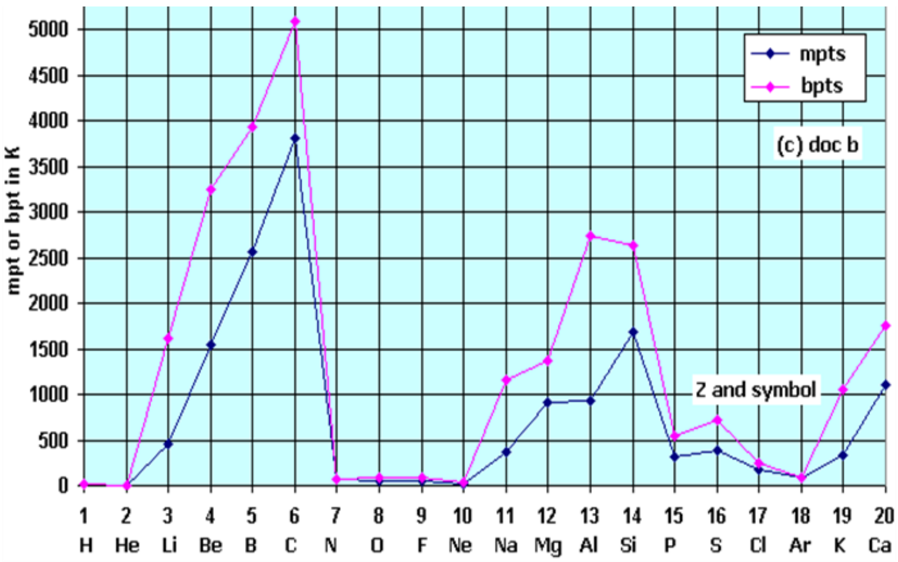
(ii) To which group do elements W belong? Explain. (2 marks)

(iii) Element Y is found to be in group IV. Write the formulae when it reacts with oxygen gas.(1mark)

(iv) Which of the two elements named in "i" above is more reactive? Explain. (2 marks)

(v) Which of the five elements above is the best conductor of electricity? Explain. (2 marks)

50. Use the figure below to answer the questions that follow

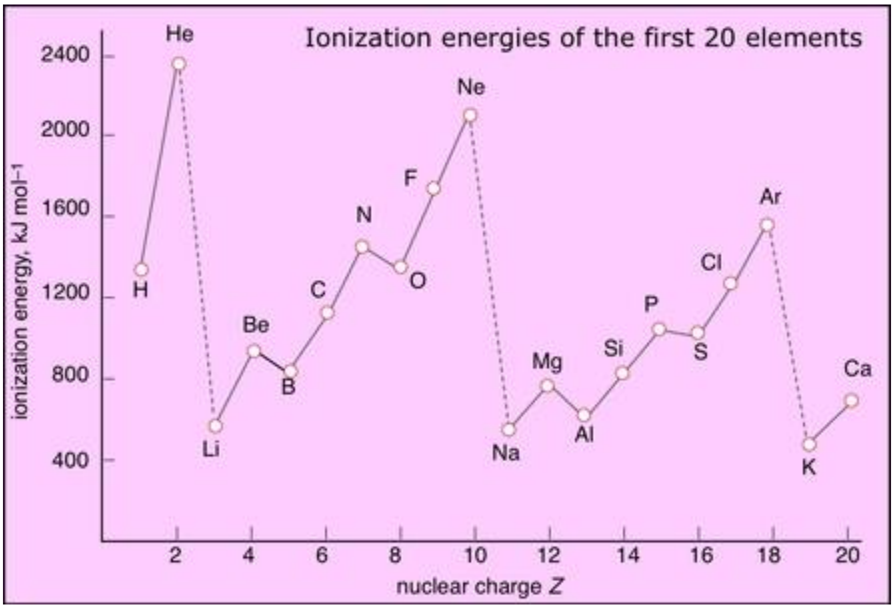


State and explain the difference in melting point of

I. Sodium and Aluminium (1 ½ marks

II. Oxygen and Sulphur (1 ½ marks

51. Use the figure below to answer the questions that follow



Explain why the first ionization energy of lithium is higher than that of potassium

I. Select the element with the highest first ionization energy

II. Select the strongest reducing agent

CHAPTER FIVE: SALTS

Specific Objectives

By the end of this topic, the learner should be able to:
a) Select and use appropriate methods of preparing particular salts
b) Explain the terms saturated solution, crystallisation, neutralisation and precipitation
c) Write ionic equations for the preparation of salts
d) State types of salts
e) Identify soluble and insoluble salts
f) Describe and explain from experimental observations the action of heat on various salts
g) State uses of some salts

SALT

Salt is a substance formed when all or part of hydrogen ions in an acid is replaced by a metal Cation or ammonium ion

TYPES OF SALTS

normal salts-they do not contain replaceable hydrogen ion eg NaCl, $MgNO_3$
acid salts-they contain replaceable hydrogen ions $NaHCO_3$,$NaHSO_4$, $NaHSO_3$,
double salts-contain two different anions or cations e.g $KAl(SO_4)_2.12H_2O$, $(NH_4)_2FeSO_4.6H_2O$
basic salts-are salts that contain hydroxyl (OH^-) ion,e.g Pb(OH)Cl, Zn(OH)Cl

Solubility of salts

- *All salts of Na, K, and NH_4^+ are soluble*
- *All nitrates (NO_3^-), acetates (CH_3COO^-), hypochlorite (OCl^-), chlorates (CLO_3^-), HCO_3^- and perchlorates (ClO_4^-) are soluble*
- *All sulphates are soluble except those of Pb^{2+}, Ba^{2+} and Ca^{2+} but $caso_4$ is sparingly soluble*
- *All chlorides are soluble except those of Pb^{2+} and Ag+ but $Pbcl_2$ soluble in hot water*
- *All metal oxides, metal hydroxides, and carbonates are insoluble except those of Na, K, and NH_4^+*

Table of solubility of salts

Soluble salts	Insoluble salts
All nitrate(V)salts	
All sulphate(VI)/SO_4^{2-} salts **except**	Barium sulphate/$BaSO_4$ Calcium sulphate/$CaSO_4$ Lead (II) sulphate/$PbSO_4$ (**SULEBACA**)
All Sulphite(IV)/SO_3^{2-} salts **except**	Barium sulphite/$BaSO_3$ Calcium sulphite/$CaSO_3$ Lead sulphite/$PbSO_3$ (**SULEBACA**)
All chlorides/Cl^- **except**	Silver chloride/AgCl Lead (II)chloride/$PbCl_2$(dissolves in hot water) **(CLESI)**
All phosphate(V)/PO_4^{3-}	
All sodium,potassium and ammonium salts	
All hydrogen carbonates/HCO_3^-	
Sodium carbonate/Na_2CO_3, Potassium carbonate/ K_2CO_3, Ammonium carbonate $(NH_4)_2CO_3$	← **except** All carbonates
All alkalis(KOH,NaOH, NH_4OH)	← **except** All bases

METHODS OF PREPARING SALTS

The method used to prepare a salt depends on whether is soluble or insoluble

METHODS OF PREPARING SOLUBLE SALTS

$i)\ Metal\ +\ acid \rightarrow salt\ +\ Hydrogen$

$Mg_{(s)} + HCl_{(aq)} \rightarrow MgCl_{2\ (aq)} + H_{2(g)}$

$ii)Acid\ + alkali\ \rightarrow salt\ +\ water$

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

$iii)Acid\ +\ insoluble\ base\ \longrightarrow salt\ +\ water$

$CuO_{(s)} + H_2SO_{4(aq)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)}$

$iv)Acid\ +\ carbonate\ \longrightarrow\ salt\ +\ water\ +\ carbon(IV)\ oxide$

$MgCO_{3(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

$v)Acid\ +\ hydrogen\ carbonate \rightarrow salt\ +\ water\ +\ carbon(IV)\ oxide$

$NaHCO_{3\ (s)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

$vi)Direct\ synthesis$

$eg,\ 2K_{(S)} + Cl_{2(g)} \longrightarrow 2\ KCl_{(s)}$

1. PREPARATION OF ZINC SULPHATE

Method; metal + acid

Reagents; zinc metal, dilute sulphuric(VI) acid

Methodology/Procedure

Measure about 20cm^3 of dilute sulphuric(VI) acid and transfer it to a clean beaker, add excess zinc powder little by little while stirring until effervescence stops, filter and heat /evaporate the filtrate to saturati**on, allow it to cool for crystals of $ZnSO_4$ to grow**

Points to note;

- Excess zinc is added to make sure all the acid has reacted
- Effervescence is due to production of hydrogen gas(***the gas can be confirmed by lowering a burning splint into a gas jar full of the gas,it is put off with a pop sound***)
- The reaction is complete when effervescence stops
- Filtration is important to remove the unreacted zinc
- Equation for the reaction

$Zn_{(s)} + H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + H_{2(g)}$

Other salts that can be prepared using the same method, $ZnCl_2$,$MgSO_4$, $CaCl_2$,$MgCl_2$

Complete the equations below

$Mg_{(s)} + H_2SO_{4(aq)} \longrightarrow$

$Zn_{(s)} + H_2SO_{4(aq)} \longrightarrow$

2. PREPARATION OF COPPER (II) SULPHATE

Method; acid + insoluble base

Reagents; dilute sulphuric (VI) acid and Copper (II) oxide

Method

Measure about 25cm^3 of dilute sulphuric (VI) and transfer it into a clean beaker, warm the acid. Add excess Copper (II) oxide little by little while stirring until no more of it can dissolve. Filter and evaporate the filtrate in an evaporating dish to saturation and cool it to allow crystals of $CuSO_4$ to grow.

Observation made

Black soild dissolves
Blue solution is formed

Points to note;

- the reaction is complete when no more copper(II) oxide can dissolve
- Warming is necessary as increases the kinetic energy of reacting particles hence increases chances of fruitful collisions hence increase rate of reaction.
- filtration is important to remove the unreacted copper(II) oxide
- copper(II) sulphate cannot be prepared using copper metal and dilute sulphuric(VI) acid because copper is below hydrogen in the reactivity series hence cant displace hydrogen.

  Equation for the reaction

  $$CuO_{(s)} + H_2SO_{4(aq)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)}$$

  Other salts that can be prepared through the above method

  Copper(II) chloride, Calcium chloride, lead(II) nitrate

$$MgO_{(s)} + H_2SO_{4(aq)} \longrightarrow$$

$$ZnO_{(s)} + H_2SO_{4(aq)} \longrightarrow$$

PREPARATION <u>OF SODIUM CHLORIDE</u>

Method; acid + alkali

Reagents- Sodium hydroxide and Hydrochloric acid

Procedure

Fill the burette with dilute hydrochloric acid, measure about 20cm^3 of dilute sodium hydroxide and transfer into a clean conical flask. Add 2-3 drops of phenolphthalein indicator to the alkali and run the acid from the burette into the alkali while shaking until there is a permanent colour change. Repeat the experiment without using the indicator but add the same amount of acid. Heat to evaporate the resulting solution to saturation and allow it to cool to obtain crystals of the salt.

Points to note;

- The reaction is complete when there is a permanent colour change
- Purpose of using the indicator to determine the end point
- The experiment is repeated without using the indicator to avoid contaminating the salt

Equation to the reaction

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

- Sodium chloride cannot be prepared using sodium metal and dilute hydrochloric acid because the reaction is explosive hence can cause accidents
- Other salts that can be prepred in the same method, complete the equations below

$$NaOH_{(aq)} + HNO_{3(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)}$$

$$KOH_{(aq)} + HNO_{3(aq)} \longrightarrow$$

$$KOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow$$

$$KOH_{(aq)} + HCl_{(aq)} \longrightarrow$$

PREPARATION OF CALCIUM CHLORIDE

Method; Acid + carbonate

Procedure

Measure about 20cm^3 of dilute hydrochloric acid and transfer it to a clean beaker, add calcium carbonate little by little until effervescence stops, filter and heat to evaporate the filtrate to saturation and allow it to cool for crystals to grow.

Points to note

- *The reaction* is complete when effervescence stops and no more carbonate can dissolve
- *$CaSO_4$ cannot be prepared using calcium carbonate and dilute sulphuric because of formation of insoluble coat of calcium sulphate which prevents further reaction between the acid and the carbonate*

Equation for the reaction

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$
$$ZnCO_{3(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$
$$CuCO_{3(s)} + 2HCl_{(aq)} \rightarrow CuCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

$PbSO_4$ and $PbCl_2$ cannot be prepared from reaction of lead carbonate with silphuric (VI) acid and hydrochloric acid due to formation of insoluble $PbSO_4$ and $PbCl_2$ that coats the carbonate preventing further reaction.
NB: preparation of salts from acids and hydrogen carbonates follow the above procedure. Only hydrogen carbonates of Na, K, Ca, Mg and NH_4^+ exist

PREPARATION OF SALTS BY DIRECT SYNTHESIS

This is preparation of salts by direct combination of elements

When a metal **burn** in a gas jar containing a non metal , the two directly combine to form a salt. e.g.

Describe how potassium chloride can be prepared by direct synthesis

cut a small piece of potassium metal, place it in a deflagrating spoon burn it briefly and lower it into a gas jar of chlorine ,it will react forming potassium chloride,

Describe how Alumnium chloride can be prepared by direct synthesis

Pass dry chlorine gas over heated aluminium metal in a combustion tube ,it will react forming Aluminium chloride

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$
$$2K_{(s)} + Cl_{2(g)} \longrightarrow 2KCl_{(s)}$$

PREPARATION OF INSOLUBLE SALTS

Insoluble salts can be prepared by reacting **two** suitable **soluble** salts to form **one soluble** and **one insoluble**. This is called **double decomposition** or **precipitation**. The mixture is filtered and the **residue** is washed with distilled water then dried

General procedure for preparing insoluble salts

The first step is obtain **two soluble salts**, theN mix the two salts to precipitate the insoluble salt, filter ,wash the residue **with distilled water to remove traces of soluble salt**, dry the residue between filter papers

$$\mathbf{Cu}SO_{4(aq)} + Na_2\mathbf{CO}_{\mathbf{3}(aq)} \longrightarrow \mathbf{CuCO}_{\mathbf{3(s)}} + Na_2SO_{4(aq)}$$
$$\mathbf{Ba}Cl_{2(aq)} + K_2\mathbf{SO}_{\mathbf{4}(aq)} \longrightarrow \mathbf{BaSO}_{\mathbf{4(s)}} + 2KCl_{(aq)}$$
$$\mathbf{Pb}(NO_3)_{2(aq)} + K_2\mathbf{SO}_{\mathbf{4}(aq)} \longrightarrow \mathbf{PbSO}_{\mathbf{4(s)}} + 2KNO_{3(aq)}$$

Points to note when answering questions on insoluble salts in the exam set up. Whereby the stater reagent is either a metal or metal oxide.

- ✓ First convert the metal/metal oxide into a soluble Salt by reacting excess metal/metal oxide with dilute nitric acid to obtain metal nitrate(all nitrates are soluble) ,then to get insoluble salt/hydroxide add a salt of potassium or sodium (because all salts of sodium and potassium are soluble ,then filter and wash the residue with distilled water and dry it between filter papers
- ✓ When the starter reagent is metal below hydrogen in the reactivity series like Copper metal you first burn the copper in air to form copper (II) oxide then react the oxide with a dilute acid.

eg lead (II) sulphate starting with lead

Add excess lead metal to dilute Nitric (V) acid to form $Pb(NO_3)_2$, filter to remove uncreated lead and $Pb(NO_3)_2$ as filtrate, add Na_2SO_4 solution to the filtrate to precipitate $PbSO_4$,filter to obtain $PbSO_4$ as residue and wash it with distilled water and dry it between filter papers .

NOTES ON PREPARATION OF DOUBLE SALTS

Double salt are prepared by mixing two soluble salts which react to form the double salt ,usually a complex salt
Preparation of ammonium iron (II) sulphate; this salt is prepared by mixing ammonium sulphate solution with iron (II) sulphate solution, then you heat to evaporate the resulting solution to saturation and allowing it to cool to form crystals then filter and dry the crystal between filter papers

In a normal examination set up the starter reagent determines the procedure to use e.g

Describe how a solid sample of the double salt .Ammonium iron (II) sulphate can be prepared using the following reagents, aqueous ammonia, and sulphuric (VI) acid and iron metal

add iron metal to dilute sulphuric (VI) acid to form iron (II) sulphate, add aquoes ammonia to sulphuric (VI) acid to form Ammonium sulphate;mix the two solutions of iron (II) sulphate and Ammonium sulphate to form a solution of Ammonium iron (II) sulphate; heat to evaporate the solution until crystallization starts to form, filter to obtain the double salt.

Properties of salts

Deliquescence , Hygroscopy and **Efflorescence**.

a) **Hygroscopic salts /compounds** are those that absorb moisture /water vapour from the atmosphere but do not form a solution e.g. copper(II)sulphate, anhydrous cobalt(II)chloride, potassium nitrate and common table salt.
NB pure sodium chloride is not hygroscorpic but **common salt contain $MgCl_2$ which give it the hygroscorpic properties as it absorbs water vapour from the atmosphere making it damp,**

b) **Deliquescent salts /compounds** are those that absorb moisture/water vapour from the atmosphere and form a solution e.g. Sodium nitrate,Calcium chloride, Sodium hydroxide, Iron(II)chloride, Magnesium chloride.

c) **Efflorescent salts/compounds** are those that lose their water of crystallization to the atmosphere e.g. sodium carbonate decahydrate, Iron(II)sulphate. heptahydrate, sodium sulphate decahydrate.

d) Some salts contain water of crystallization.They are hydrated.Others do not contain water of crystallization. They are anhydrous.

Table showing some hydrated salts.

Name of hydrated salt	**Chemical formula**
Copper(II)sulphate(VI)pentahydrate	$CuSO_4.5H_2O$
Aluminium sulphate hexahydrate	$Al_2 (SO_4)_3.6H_2O$
Zinc(II)sulphate heptahydrate	$ZnSO_4.7H_2O$
Iron (II) sulphate heptahydrate	$FeSO_4.7H_2O$

Table of some double salts

Name of double salts	Chemical formula
Trona (sodium sesquicarbonate dehydrate)	$\mathbf{Na}_2CO_3 \bullet \mathbf{NaHCO_3} \bullet 2H_2O$
Ammonium iron (II) Sulphate	$(NH_4)_2 Fe(SO_4)_2 .6H_2O$
Ammonium Aluminium Sulphate	$(\mathrm{NH_4})\mathrm{Al}(\mathrm{SO_4})_2 \cdot 12\mathrm{H_2O}$

Effects of heat on salts

a) Effect of heat on chlorides

All chlorides have very high melting and boiling points and therefore are not affected by laboratory heating except ammonium chloride. Ammonium chloride **undergoes** thermal **dissociation** into **ammonia** and **hydrogen chloride** gases on strong heating.

$$NH_4Cl_{(g)} \rightleftharpoons NH_{3(g)} + HCl_{(g)}$$

b) Effect of heat on nitrates

i. Potassium nitrate/KNO_3 and sodium nitrate/$NaNO_3$ decompose on heating to form Potassium nitrate/KNO_2 and sodium nitrate/$NaNO_2$ and producing Oxygen gas in each case.

$$2KNO_{\mathbf{3}(s)} \xrightarrow{heat} 2KNO_{2(s)} + O_{2(g)}$$

$$2NaNO_{\mathbf{3}(s)} \xrightarrow{heat} 2NaNO_{2(s)} + O_{2(g)}$$

ii. Heavy metal nitrates(V) salts decompose on heating to form the oxide and a mixture of brown acidic nitrogen(IV)oxide and oxygen gases. e.g.

$$2Ca(NO_3)_{2(s)} \longrightarrow 2CaO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

$$2Mg(NO_3)_{2(s)} \longrightarrow 2MgO_{(s}) + 4NO_{2(g)} + O_{2(g)}$$

iii. Silver(I)nitrate and Mercury (II) nitrate are lowest in the reactivity series. They decompose on heating to form the **metal** (silver and mercury)and the Nitrogen(IV)oxide and oxygen gas. i.e.

$$2AgNO_{3(s)} \xrightarrow{heat} 2Ag_{(s)} + 2NO_{2(g)} + O_{2(g)}$$

$$2Hg(NO_3)_{2\ (s)} \xrightarrow{heat} 2Hg_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

Ammonium nitrate and Ammonium nitrate decompose on heating to Nitrogen(I)oxide(relights/rekindles glowing splint) and nitrogen gas respectively.Water is also formed.i.e.

$$NH_4NO_{3(s)} \rightarrow N_2O_{(g)} + H_2O_{(l)}$$

$$NH_4NO_{2(s)} \rightarrow N_{2(g)} + H_2O_{(l)}$$

NB:The ease of decomposition of nitrates increases down the reactivity series.

c) Effect of heat on sulphates

Most metal sulphates are very stable and don't decompose on heating .Only Iron(II)sulphate, Iron(III)sulphate and copper (II) sulphateand Zinc sulphate decompose on heating. They form the **oxide**, and produce highly acidic fumes of acidic **sulphur (VI)oxide** gas.

$$2FeSO_4.7H_2O_{(s)} \rightarrow Fe_2O_{3(s)} + SO_{3(g)} + SO_{2(g)} + 14H_2O_{(l)}$$

$$Fe_2(SO_4)_{3(s)} \rightarrow Fe_2O_{3(s)} + SO_{3(g)}$$

$$CuSO_4.5H_2O_{(s)} \rightarrow CuO_{(s)} + SO_{3(g)} + 5H_2O_{(l)}$$

$$2\,ZnSO_4.7H_2O_{(s)} + 2ZnO_{(s)} + 2SO_{3(g)} + 14H_2O_{(l)}$$

Effect of heat on carbonates and hydrogen carbonate.

i. Sodium carbonate and potassium carbonate **do not decompose** on heating.

ii. Heavy metal nitrate salts decompose on heating to form the **oxide** and produce **carbon(IV)oxide** gas. Carbon (IV)oxide gas forms a white precipitate when bubbled in lime water. The white precipitate dissolves if the gas is in excess. e.g.

$$CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$$

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

$$PbCO_{3(s)} \rightarrow PbO_{(s)} + CO_{2(g)}$$

iii. Sodium hydrogen carbonate (IV) and Potassium hydrogen carbonate (IV)decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i.e.

$$2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$$

$$2KHCO_{3(s)} \rightarrow K_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$$

iv. Calcium hydrogen carbonate (IV) and Magnesium hydrogen carbonate(IV) decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i. e.

$$Ca(HCO_3)_{2(\mathbf{aq})} \longrightarrow CaCO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$$

$$Mg(HCO_3)_{2(\mathbf{aq})} \longrightarrow MgCO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$$

NB: $Ca(HCO_3)_{2(\mathbf{aq})}$ and $Mg(HCO_3)_{2(\mathbf{aq})}$ only exist in solution form.

Summary of action of heat on salts

Metal	Action of heat on Nitrates	Action of heat on carbonates
K Na	Decompose to metal nitrite and oxygen gas $2NaNO_{3(s)} \xrightarrow{heat} 2NaNO_{2(s)} + O_{2(g)}$	No effect
Ca Mg Al Zn Fe Pb Cu	Decompose to form metal oxide ,nitrogen(IV) oxide and oxygen gas e.g. $2Pb(NO_3)_{2(s)} \xrightarrow{heat} 2PbO_{(s)} + 4NO_{2(g)} + O_{2(g)}$ $2Cu(NO_3)_{2(s} \xrightarrow{heat} 2CuO_{(s)} + 4NO_{2(g)} + O_{2(g)}$	Decompose to form metal oxide and **carbon(IV)oxide** gas. eg $PbCO_{3(s)} \xrightarrow{heat} PbO_{(s)} + CO_{2(g)}$ $FeCO_{3(s)} \xrightarrow{heat} FeO_{(s)} + CO_{2(g)}$
Hg		

Ag	\Decompose to free metal, Nitrogen (IV) oxide and oxygen gas e.g. $2Hg(NO_3)_{2(s)} \xrightarrow{heat} 2Hg_{(s)} + 4NO_{2(g)} + O_{2(g)}$ $2AgNO_{3(s)} \xrightarrow{heat} 2Ag_{(s)} + 2NO_{2(g)} + O_{2(g)}$	Silver carbonate decompose to silver and CO_2 $Ag_2CO_{3(s)} \xrightarrow{heat} Ag_2O_{(s)} + CO_{2(g)}$

Decomposition of ammonium salts

$$(NH_4)_2CO_{3(s)} \xrightarrow{heat} NH_{3(g)} + CO_{2(g)} + H_2O_{(l)}$$

$$NH_4NO_{2(s)} \xrightarrow{heat} N_{2(g)} + H_2O_{(l)}$$

$$NH_4NO_{3(s)} \xrightarrow{heat} N_2O + H_2O_{(l)}$$

$$NH_4HCO_{3(s)} \xrightarrow{heat} NH_{3(g)} + CO_{2(g)} + H_2O_{(l)}$$

SPECIAL ANALYSIS OF SALTS

These are special notes prepared in order to demystify the concept of salt as tested in the KCSE exam ,it is my hope students will find them helpful and will help the to unravel the mystery of salts and consequently improve in their grades in chemistry because chapter of salts is a core topic and is always tested in KCSE. However these notes should be used together with the teachers notes and practical manual notes in demystifying chemistry practical guide book. Also it is important to that it will be almost impossible to answer any question on salts without knowledge on solubility of salts.

Questions on salts will be based on three main areas:

- **Preparation**
- **Separation**
- **action of heat and other properties eg, hygroscopy ,efflorescence ,and deliquescence**

QUESTIONS ON PREPARATION

1. *To answer question on preparation of salts you need to have solubility of salts on your nerves if not on your fingertips, but most questions test on preparation of insoluble salts of which the concept of precipitation must be used. Also when reacting acid with a base, metal or carbonate you must make sure all the acid is used up by adding excess metal or excess carbonate to the acid and always filter to remove the unreacted solid. Do not add excess acid because removing the excess acid is difficult.*
2. *When preparing salts, the method used is determined by starter reagent and if it is soluble, insoluble or a double salt.*
3. *To prepare a soluble salt whereby the starter reagent is metal/metal oxide/metal hydroxide or metal carbonate. Add excess of the metal/metal oxide or metal carbonate to dilute acid, filter heat the filtrate to saturation and cool to crystalize.*
4. *For soluble salt preparation, add excess reagent to dilute acid, filter, evaporate the filtrate to saturation, cool it to crystallize, filter*
5. *For insoluble salt the method must involve precipitation by reacting two soluble salts, for instance: to get the soluble salt whereby the starter reagent is metal/metal oxide/metal hydroxide or metal carbonate. Add excess of the metal/metal oxide or metal carbonate to dilute nitric (V) acid (because all nitrates are soluble), filter then add a soluble salt of sodium or potassium (because all salts of potassium and sodium are soluble) to precipitate the insoluble salt. Filter, wash the residue with distilled water and dry it between filter papers.*
6. *In case the salt to be prepared is a salt of copper or silver (metals below hydrogen in the reactivity series) first burn the metal in air to form metal oxide (because metals below hydrogen in reactivity series don't react with dilute acids), then add excess metal oxide to dilute acid, filter and add soluble salt of sodium or potassium. filter wash the residue with distilled water and dry it between filter papers.*

7. *In preparation of insoluble salt whereby you are given the solid soluble salt to use, first dissolve the salt in water to obtain its solution*

8. NB when considering insoluble salts remember oxides like CuO, and other metal oxides are insoluble except those of Na and

Sample questions and solutions

1. **Starting with PbO explain how you can prepare a solid sample of $PbSO_4$**

For the above question refer to guide number 4 ,first convert PbO to $Pb(NO_3)_2$,then react with a soluble sulphate of sodium or potassium

Add excess PbO to dilute Nitric (V) acid to form $Pb(NO_3)_2$, filter to remove uncreated PbO as residue and $Pb(NO_3)_2$ as filtrate, add Na_2SO_4 solution to the filtrate to precipitate $PbSO_4$,filter to obtain $PbSO_4$ as residue and wash it with distilled water and dry it between filter papers .

- **Common mistake students make is that the start by reacting lead metal with dilute Sulphuric (VI) acid to form $PbSO_4$ and reaction by between lead metal and the acid immediately stops due to formation of insoluble $PbSO_4$ which coats the metal preventing further reaction between the acid and metal, the candidate loses all the marks**

2. **Starting with copper metal describe how you can prepare a solid sample of $CuCO_3$, refer to guide number 5.**

Heat copper in air to form CuO, add excess CuO to dilute HCl to form $CuCl_2$, filter to remove unreacted CuO and $CuCl_2$ as filtrate, add aqueous Na_2CO_3 to the filtrate to precipitate $CuCO_3$ filter to obtain $CuCO_3$ as residue wash it with distilled water and dry it between filter papers.

- **Common mistake students make is that they start by reacting copper metal with dilute acids and copper being below hydrogen in the reactivity series does not react with dilute acids and candidates loses all the marks, such candidates are called as non-starters**

3. **Starting with sodium metal explain how a solid sample of $NaHCO_3$ can be prepared**

Cut a small piece of sodium metal and place it in a given amount of distilled water in a beaker,it will react to form sodium hydroxide solution Bubble excess carbon (IV) oxide in sodium hydroxide solution to form sodium hydrogen carbonate. Heat the sodium hydrogen Carbonate solution to saturation and allow it cool for crystals of $NaHCO_3$ to grow. Filter and dry the wet crystals between filter papers

4. **Starting with sodium metal explain how a solid sample of $NaHSO_3$ can be prepared**

Cut a small piece of sodium metal and place it in a given amount of distilled water in a beaker,it will react to form sodium hydroxide solution. Bubble excess sulphur (IV) oxide in sodium hydroxide solution to form sodium hydrogen sulphate solution. Heat to evaporate the resulting solution to saturation and allow it cool for crystals of $NaHSO_3$ to grow .filter and dry the wet crystals between filter papers

5. **Describe how a sample of sodium chloride can be prepared in the laboratory by direct Synthesis.**

Cut a piece of Sodium: metal, place it on a deflagrating spoon, heat it briefly then lower it : into a gas jar of chlorine . It will continue burning forming Sodium Chloride. :

In preparation of insoluble salt whereby you are given the reagents to use and one of the reagents to be used is a solid soluble salt. First dissolve the soluble salt in distilled water to obtain a solution. See the example below.

6. **Describe how the following reagents can be used to prepare Lead sulphate, solid potassium sulphate, solid lead carbonate, dilute nitric acid and distilled water**.

dissolve K_2SO_4 in distilled water to obtain K_2SO_4 solution, Add excess $PbCO_3$ to dilute HNO_3 and stir, filter to obtain $Pb(NO_3)_2$ as filtrate,. Mix $Pb(NO_3)_2$ solution with K_2SO_4 solution to precipitate $PbSO_4$,filter and wash the residue with distilled water and dry it between filter papers.

7. **You are provided with the following:- solid lead (II) nitrate, magnesium oxide powder, dilute sulphuric (VI)acid and distilled water. Describe how you can prepare a dry. Sample of *lead (II) sulphate***

Dissolve lead (II) Nitrate crystals in a given amount of distilled water in a beaker ,add excess magnesium√ ½ oxide powder To dilute sulphuric√ ½ (VI) acid in a beaker ,filter , mix the two solutions obtained ,filter wash the residue with distilled water and Dry it between filter papers to obtain a dry sample of lead (II) sulphate.

8. ***starting with solid sodium chloride describe how a pure sample of lead (II) chloride can be prepared in the laboratory***

Dissolve sodium chloride in distilled water, add aqueous lead (II) nitrate *to the sodium chloride solution, filter, the mixture ,*Wash the residue *with distilled water ,dry the residue between filter papers.*

9. In the preparation of magnesium carbonate, magnesium was burnt in air and the product Collected. Dilute sulphuric acid was then added and the mixture filtered and cooled. Sodium carbonate was added to the filtrate and the contents filtered. The residue was then washed and dried to give a white powder.
   a) Give the name of the product formed when magnesium was burnt in air.
   Magnesium Oxide
   b) Write the chemical equation for the formation of the product
   $2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$
   c) Name the filtrate collected after sodium carbonate was added.
   Sodium sulphate
   d) Write down the chemical formula of the white powder
   $MgCO_3$
   e) Write a chemical equation for the reaction between product in (a) and the acid
   $MgO_{(s)} + H_2SO_{4(aq)} \longrightarrow M_gSO_{4(aq)} + H_2O_{(l)}$
   f) Write an ionic equation to show the formation of the white powder.
   $Mg^{2+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \longrightarrow M_gCO_{3(s}$
   g) Write an equation to show what happens when the white powder is strongly heated.
   $M_gCO_{3(g)} \longrightarrow MgO_{(g)} + CO_{2(g)}$
   h) Identify the ions present in the filtrate after addition of sodium carbonate.
   Na^+ ions and SO_4^{2-} ions
   i) What is the name given to the reaction that takes place when sodium carbonate was added to the filtrate?
   Precipitation/ double decomposition

10. **Starting with sodium oxide, describe how a sample of crystals of sodium hydrogen carbonate may be prepared**

Add water to sodium oxide to form sodium hydroxide solution. Bubble excess carbon (IV) oxide in sodium hydroxide solution to form sodium hydrogen carbonate. Heat sodium hydrogen Carbonate solution to saturation and allow it to cool to crystallize .filter and dry the crystals between filter papers

Question of salts whereby you are given the exact quantity of one the reactants to use ,in such question you must invoke the concept of mole concept to get the quantity of the other reactant, in such question you don't need to filter as the reactants react completely and none is in excess. Examples are given below

Questions on separation

Question on separation require candidates to understand properties of salt about their solubility, sublimation and other related substances like metal oxides and substances like sulphur ,iodine etc

To separate a mixture containing soluble and insoluble substance

Is such questions assume you want to separate X and Y, X is soluble Y is insoluble

The answer format…. Add distilled water to the mixture and stir, X dissolve and Y does not dissolve, filter to obtain Y as residue and solution of X as filtrate, heat the filtrate to saturation and allow it to cool for crystal to form .filter,
To answer a question on separation involving a soluble and an insoluble salt e.g

11. **Describe the process of separating sodium chloride crystals mixed with calcium carbonate powder.**

Add distlled water to the mixture, stir to dissolve sodium chloride, filter to obtain calcium carbonate as residue and sodium chloride as filtrate, wash the residue with distilled water and dry the residue of calcium carbonate between filter papers, evaporate the filtrate to obtain crystals of sodium chloride. *Filter and dry the crystals between filter papers*

NB: many students lose marks start by dissolve the mixture; This implies that the whole mixture is soluble in water and therefore contradicts the whole separation process. In this case any answer beginning with the word "dissolve' renders the whole process wrong and therefore candidate losses all the marks. This is called nonstarters

12. **Some KCl was found to be contaminated with CuO .Describe how a sample of KCl can be obtained from the mixture.**

Add distilled water to the mixture and stir, KCl dissolves while CuO does not dissolve,filter to obtain KCl solution as filtrate,evaporate the filtrate to saturation and allow it to cool for crystals to grow. filter and dry the crystals between filter papers

When separating a mixture containing soluble and insoluble salts it is important to note that lead (II) chloride is soluble in hot water hence a mixture containing lead (II) chloride and another insoluble salt can be separated by …… add hot water to the mixture to dissolve $PbCl_2$, filter while still hot .see example below

13. **Describe how a mixture of $PbCl_2$, NaCl and AgCl can be separated to obtain each of the salts**

Add cold distilled water to the mixture and stir, NaCl dissolves while $PbCl_2$ and AgCl *do not dissolve, filter to obtain NaCl solution as filtrate , heat the filtrate to dryness to obtaind NaCl solid, Add hot distilled water to the remaining mixture in a beaker and stir thoroughly, $PbCl_2$ dissolves in hot water AgCl does not dissolve, filter while still hot to get AgCl as residue and $PbCl_2$ solution as filtrate, cool the filtrate to precipitate lead (II) chloride, filter to obtain $PbCl_2$ as residue and dry it between filter papers*

14. **Given a solid sample of calcium carbonate and lead (II) Chloride ,explain how you can obtain some pure crystals of lead (II) chloride**

Add hot distilled water to the mixture in a beaker and stir thoroughly, $PbCl_2$ dissolves in hot water $CaCO_3$ does not dissolve, filter while still hot to get $PbCl_2$ as filtrate, cool the filtrate for to precipitate lead (II) chloride, filter to obtain $PbCl_2$ as residue and dry it between filter papers

15. **A sample of copper turnings was found to be contaminated with copper (II) oxide. Describe how a sample of copper metal can be separated from the mixture**

Adds excess dilute hydrochloric acid/ sulphuric (vi) acid CuO reacts with H_2SO_4 while copper metal does not react,Filter to obtain copper metal,Wash with distilled water

To separate a mixture containing substance that sublimes

To answer a question on separation involving a salt that sublimes. Always start your answer in heating the mixture in a container covered with evaporating dish with cold water, where the substance that sublimes will sublime and collect on the evaporating dish. Substances that sublime include $FeCl_3$, $AlCl_3$, , iodine ,dry ice, benzoic acid etc

Note : Ammonium chloride *undergoes thermal dissociation to form ammonia gas and hydrogen chloride gas which on cooling* , the products of heated ammonium chloride recombine together to form the original substance. This process is called ***thermal dissociation***. However, although ***Ammonium chloride*** does not sublimes when heated, the principle of sublimation can be used to separate a mixture containing ammonium chloride

16. **Given a mixture of lead (II) oxide, iodine and sodium chloride, describe how this mixture can be separated to obtain a sample of each.**

Heat the mixture in a container covered with evaporating dish with cold water iodine sublimes and is collected on the evaporating dish. Add water to the remaining mixture, stir and filter. Lead (ii) Oxide remains as residue. Heat to evaporate th filtrate to dryness to obtain sodium chloride, filter

NB. Common mistake students don't use a closed container, and also state that iodine sublimes and condenses on cooler parts. Condensing vapour implies that iodine becomes liquid which contradicts the principle of sublimation. This makes the answer wrong and the candidate loses all the marks.

17. Describe how solid ammonium chloride can be separated from a solid mixture ammonium chloride and anhydrous calcium chloride.

Heat the mixture in a container covered with a evaporating dish with water, collect NH4Cl on the pevaporating dish, $CaCl_2$ remains at the bottom of the container.

18. Describe how samples of lead(II) sulphate ,sodium chloride and Ammonium Chloride can be obtained from a mixture of the three.

Heat the mixture in a container covered with evaporating dish with water and collect NH_4Cl on the evaporating dish while NaCl and $PbSO_4$ remain behind , ,add water to the remaining mixture and stir, NaCl dissolves $PbSO_4$ doesn't filter to obtain $PbSO_4$ as residue and NaCl solution as filtrate ,heat to evaporate the filtrate to dryness to obtain solid NaCl. .

Questions to differentiate salts and other substances,

Here you need to understand reactions of salts including thermal decomposition, reacting with acids and colour of coloured ions, you also need to apply the knowledge of qualitative analysis in order to answer such questions Examples given below

19. **Without using any laboratory chemical, describe a simple laboratory experiment to distinguish Between calcium hydrogen carbonate and sodium hydrogen carbonate solutions** .

EITHER

In separate test tubes, boil about 5cm³ of each solutions, Sodium hydrogen carbonate solution remains colourless/ forms no precipitate, Calcium hydrogen carbonate solution changes from colourless to white precipitate

OR

$$2NaHCO_{3aq} \xrightarrow{heat} Na_2CO_3 + CO_{2(g)}n + H_2O_{(e)}$$
$$Ca\ (HCO_3)_{2(aq)} \xrightarrow{heat} CaCO_{3(s)} + CO_{2(g)} + H_2O_{(e)}$$

HEAT must be mentioned or implied

20. Copper (II) oxide and charcoal are black solids. How would you distinguish between them.

To separate samples of CUO and charcoal in test tubes, add dilute dilute sulphuric (VI) acid and shake black CuO black dissolves to form blue solution √ and no effect on charcoal

21. ***Describe how $NaHCO_3$ and Na_2CO_3 and be differentiated***.

Heat the two solids separately in test tubes $NaHCO_3$ evolves a gas that turns blue litmus red and colourless liquid condenses on cooler parts of test tube, Na_2CO_3 not affected by heating

22. Describe one method that can be used to distinguish between sodium sulphate and sodium hydrogen sulphate.

Test the acidity using a litmus pager. There will be no change on blue litmus when dipped into a solution of sodium sulphate (1). The blue litmus paper turns to red when dipped into a solution of sodium hydrogen sulphate (l).

OR

Add a solid carbonate to each solution. No effervescence observed when the carbonate is added to a solution of sodium sulphate. Effervescence is observed when the carbonate is added to a solution of sodium hydrogen sulphate.

23. Describe how CuO, MnO_2 and FeS can *differentiated* using dilute HCl only

Place the three substances into separate test-tubes and add dilute HCl

CuO react to form a green solution,MnO_2 reacts to form a colourless solution, and FeS reacts to form a green solution

Most students assume that all copper salts are blue in colour ,this is wrong as some copper salts like copper (II) carbonate and copper (II) chloride are green

Copr

24. ***Describe one method that can be used to distinguish between sodium hydrogen carbonate and sodium hydrogen sulphate***

Add distilled water to each of the salts , put red and blue litmus papers to each of the solutions ,there will no change on blue litmus paper ,while red litmus paper turns blue when put to the solution of sodium hydrogen carbonate,the blue litmus paper turns red and no effect on red litmus when dipped in solution of sodium hydrogen sulphate.

25. ***Describe one method that can be used to distinguish between sodium hydrogen carbonate and sodium hydrogen sulphite***

Add distilled water to each of the salts , put red and blue litmus to each of the solutions ,there will no change on blue litmus ,while red litmus turns blue when put to the solution of sodium hydrogen carbonate,the blue litmus paper turns red and no effect on red litmus when dipped in solution of sodium hydrogen sulphite.

Common mistake most students make is that they assume that acid salts all acid salts dissolve in water to form acidic solution, this is incorrect as acid salt means the salt has replaceable hydrogen ion, therefore aqueous solution of $NaHCO_3$ is alkaline not acidic, but aqueous $NaHSO_3$ and $NaHSO_4$ are acidic

26. ***Describe how you can differentiate between sodium sulphite and sodium carbonate,***

Add distilled water to each of salts ,the add acidified potassium manganate(VII) to solutions of each, potassium manganate (VII) turn from purple to colourless in sodium sulphate while the purple colour persist in sodium carbonate

27. ***Describe one method that can be used to distinguish between sodium sulphate and sodium sulphite***

Add distilled water *to each of the salts , then add acidified barium chloride/barium nitrate ,a white precipitate will be formed with sodium sulphate and no precipitate with sodium sulphite.*

28. **Describe a simple laboratory experiment that can be used to distinguish between sodium sulphide and sodium carbonate.(3mk)**

Add dilute hydrochloric√/HCl//sulphuric(VI)acid/H_2SO_4 to each separately.
With sodium sulphide a colourless gas with the smell of rotten eggs is evolved √.
With sodium carbonate a colourless odourless gas is evolved √.

Question on effect on salts on exposure to air,

Here the candidates need to understand on meaning of **efflorescence**, **deliquescence** and **hygroscopy** and know examples of salts and substances that undergo the aforementioned processes

29. **Explain why anhydrous NaOH pellets when exposed to air ,first turn into a colourless solution and finally into a white powder**

NaOH is deliquescent and absorb water from the atmosphere to form NaOH solution, the NaOH solution react with CO_2 to form hydrated sodium carbonate which is efflorescent and loses water to form anhydrous Na_2CO_3

30. **Pellets of NaOH and anhydrous $CuSO_4$ were put in different petri dishes and left in the open for two hours , explain the observation that was made.**

The petri dish containing NaOH had absorbed so much water from the atmosphere to form a solution ,petri dish containing $CuSO_4$ becomes damp and colour changed from white to blue.

Reasoning NaOH is deliquescent while $CuSO_4$ is hygroscorpic

31. Explain the observations made when crystals of sodium carbonate decahydrate are left exposed to the atmosphere for two days

Crystals turn to a white powder. The salt is efflorescent hence it loses its water of crystallization forming a powder

32. Give the name of each of the processes described below which takes place when salts are exposed to air for sometime
    a) Anhydrous copper sulphate becomes damp ***Hydroscopy***
    b) Magnesium chloride forms an aqueous solution **Deliquescence**
    c) Fresh crystals of sodium carbonate, $Na_2CO_3.10H_2O$ become covered with white powder of formula $Na_2CO_3.H_2O$ ***Efflorescence***

NOTE

I. **Efflorescent substances** (e.g. sodium carbonate decahydrate, copper (II) sulphate pentahydrate) lose some or all of their water of crystallization when exposed to the air.

II. **Hygroscopic substances** (e.g. concentrated sulphuric acid, calcium oxide) absorb water vapour/moisture from the atmosphere but do not dissolve in it.

III. **deliquescence,** the process by which a substance absorbs water vapour from the atmosphere until it dissolves in the absorbed water and forms a solution. Eg of deliquescent substances include calcium chloride,$ZnCl_2$, $FeCl_3$, $MgCl_2$.

Common mistake …many students write **absorb water** from the atmosphere **instead of water vapour/moisture**, atmosphere contains water vapour not water

33. **Sodium chloride is not hygroscorpic but common salt when left to the atmosphere becomes damp, explain**,

common salt contain $MgCl_2$ which give it the hygroscorpic properties as it absorbs water vapour from the atmosphere making it damp,

Question on action of heat on salt

You must understand the action of heat on carbonates and nitrates ,some sulphates like $ZnSO_4$, $CuSO_4$ $FeSO_4$, also note action of heat on $(NH4)_2CO_3$, NH_4NO_3,.$(NH_4)_2SO_4$ and NH_4Cl

NB NH_4Cl undergoes l decomposition on heating to NH_3 and HCl gases,NH_3 is lighter and diffuses faster Hence the litmus first turns from red to blue, and later the HC l reaches the upper end of test-tube and turns blue litmus red

34. A form two student was asked to prepare a sample of copper (II) sulphate crystals using the procedure below.
    - Measure 100cm^3 of 2M sulphuric acid then warm. Add excess copper (II) oxide powder. Filter the resulting mixture.
    - Heat the filtrate and leave it overnight.

a) Why was the acid heated before the start of the reaction?

To increase the rate of reaction

b) Why was excess copper(II) oxide used (1 mark)

To make sure that all the acid has reacted

c) What was observed when copper (II) oxide was added to the warm acid? (2 marks)

A blue solution is formed,black CuO dissolves

d) Write an equation for the reaction that took place in (c) above (1 mark)

$$H_2SO_{4(aq)} + CuO_{(s)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)}$$

e) Give reasons for carrying out the following processes

i. Filtration of the mixture (1 mark)

To eliminate excess//unreacted CuO

ii. Heating the filtrate and leaving it overnight (2 marks)

To drive out//evaporate some H_2O and for the saturated solution to cool and form crystals.

f) Explain how dry crystals of copper (II) sulphate are finally obtained. (1 mark)

Wet crystals were dried between filter papers.

g) State and explain the observations that would be made when concentrated sulphuric (VI) acid is added to the crystals formed in (f) above in a test tube.(2 marks)

Blue crystals turn to white powder

Exp; Conc H2SO4 is a dehydrates $CuSO_4$ x H_20 crystals to anhydrous $CuSO_4$.

h) Write the formula of the complex ion formed with excess ammonia solution is added to copper (II) sulphate solution.

$$\left[Cu\left(NH_3\right)_4\right]^{2+}_{(aq)}$$

Explain why it would not be possible to prepare copper sulphate salt by reaction of dilute sulphuric (VI) acid with copper metal?

Copper is below hydrogen in the reactivity series and so cannot displace it from the acid

Reversed salt preparation. Is such questions you might be required to describe how to obtain a metal oxide or a metal from a salt solution , you can heat the salt solution to dryness, then heat the dry salt to decompose to metal oxide, you can also add a soluble carbonate then heat the carbonate strongly . To obtain the metal oxide ,to obtain a metal then you can reduce the oxide(while hot) with carbon (II) oxide metal to obtain metal.

35. ***Starting with Zinc sulphate solution ,describe how a sample of zinc oxide can be obtained from a solution of Zinc (II) Sulphate***

Add soluble Sodium carbonate /add NaOH
Filter out the Zinc carbonate /filter the $Zn(OH)_2$ */ heat strongly the* $Zn(OH)_2/ZnCO_3$ *to decompose it to form ZnO*
Or heat to evaporate the water to obtain $ZnSO_4$ *solid*
Heat $ZnSO_4$ *solid strongly to decompose to form ZnO*

36. ***Describe how you can obtain lead metal from lead(II) carbonate***

Heat the $PbCO_3$ *strongly in a crucible to decompose to PbO, pass hydrogen gas/CO over the hot PbO to reduce it to lead metal*

37. ***Describe how you can obtain MgO from Magnesium Sulphate solution***

Add aqueous Sodium carbonate to aqueous Magnesium Sulphate to precipitate $MgCO_3$. Filter out the magnesium carbonate , heat strongly the $MgCO_3$ in a crucible to decompose it to form MgO

38. **Starting with solid Aluminium Sulphate, describe how a solid sample of Aluminium hydroxide could be prepared. (3mks)**

Add distilled water to Aluminium Sulphate to dissolve. √½
Add aqueous ammonia hydroxide √ to Aluminium Sulphate solution, to precipitate Aluminum hydroxide, Filter. √½,Wash the residue with distilled water. Dry it between filter papers √½

When preparing $Al(OH)_3$, or $Pb(OH)_2$ aqueous NaOH can not be used because the amphoteric $Al(OH)_3$, or $Pb(OH)_2$ will dissolve in excess, hence aqueous ammonia is the suitable alkali to use.

39. **Starting with lead metal, describe how a solid sample of lead (II) hydroxide could be prepared. (3mks**

Add excess lead metal to dilute Nitric (V) acid to form $Pb(NO_3)_2$*, filter ,* Add aqueous ammonia hydroxide solution√ to
The filtrate, to precipitate lead (II) hydroxide, Filter. √½, Wash the residue with distilled water. Dry it between filter papers √½

40. **Describe how a solid sample of Zinc (II) carbonate can be prepared starting with zinc oxide(3mk)**

Add excess zinc oxide to dilute HCl, H_2SO_4 */* HNO_3*√., Filter√½ to remove excess zinc oxide., Add* Na_2CO_3 *,* K_2CO_3 *solution to the filtrate√½.Filter√½.Wash the residue with distilled water. Dry between filter papers√½.*

PREPARATION OF DOUBLE SALTS

Double salt are prepared by mixing two soluble salts which react to form the double salt, usually a complex salt preparation of ammonium iron (II) sulphate; this salt is prepared by mixing ammonium sulphate solution with iron (II) sulphate solution, then you evaporate the resulting solution to saturation and allowing it to cool to form crystals then filter and dry the crystal between filter papers. **In a normal examination set up the starter reagent determines the procedure to use e.g**

41. **Describe how a solid sample of the double salt. Ammonium iron (II) Sulphate can be prepared using the following reagents, aqueous ammonia, sulphuric (VI) acid and iron metal**
Add iron metal to dilute sulphuric (VI) acid to form iron (II) Sulphate, add aqueous ammonia to sulphuric (VI) acid to form Ammonium Sulphate; mix the two solutions a of iron (II) Sulphate and Ammonium Sulphate to form a solution of Ammonium iron (II) Sulphate; heat to evaporation the solution until crystallization starts, filter to obtain the double salt.

42. **Starting with calcium and dilute nitric (V) acid and ammonia describe how a solid sample of calcium ammonium nitrate can be prepared**
Add calcium metal to dilute nitric (V) acid to form calcium nitrate solution, react aqueous ammonia with nitric (V) acid to form Ammonium nitrate solution; mix the two solutions of calcium nitrate and Ammonium nitrate to form a solution of calcium ammonium nitrate; evaporation the solution until crystallization starts, filter to obtain the double salt

43. **Describe how a solid sample of the double salt. Ammonium Aluminium Sulphate can be prepared using the following reagents, aqueous ammonia, sulphuric (VI) acid and Aluminium metal:**
Add *Aluminium metal with dilute sulphuric (vi) acid to form* Aluminium *Sulphate, add aqueous ammonia with sulphuric (VI) acid to form ammonium Sulphate; mix the two solutions of* Aluminium *Sulphate and ammonium Sulphate to form a solution of* ammonium Aluminium Sulphate; heat to evaporation *the solution until crystallization starts, filter to obtain the double salt.*

44. **Hydrated cobalt(II)chloride exist as pink crystals and anhydrous cobalt(II)chloride is a blue powder. Describe a laboratory experiment that can be used to show that the action of heat on hydrated cobalt(II)chloride is a reversible reaction. (3mk)**

Heat the pink hydrated salt in a sealed container. √
Pink substance changes to blue. √ ½
Collect any vapour and cool it to condense. √ ½
Add the condensed liquid to the blue solid. √ ½
It turns back to pink. √ ½

CHAPTER SIX: **EFFECT OF ELECTRIC CURRENT OMN SUBSTANCES**

DEFINITIONS

- **Electrolysis** –refers to decomposition of an electrolyte into its constituent ions by passing an electric current through it
- **Electrolyte –**a solution or a melt of a compound that allows an electric current to pass through it and it decomposed by it. Some electrolytes are **weak** while others are **strong**.
- **Strong electrolytes** are those that are fully ionized/dissociated into (many) ions. Common strong electrolytes include:
  i. all **mineral** acids : dilute HCl ,HNO_3, H_2SO_4
  ii. all strong **alkalis**/sodium hydroxide/potassium hydroxide.
- **Weak electrolytes** are those that are partially/partly ionized/dissociated into (few) ions. Common weak electrolytes include:
  i. All **organic** acids
  ii. All **bases** except sodium hydroxide/potassium hydroxide.
  iii. Water
- **Non electrolyte –**solutions or melts that do not conduct an electric current. Non-electrolytes are those compounds /substances that exist as molecules and thus cannot ionize/dissociate into (any) ions. Common non-electrolytes include:
  i. Most organic solvents (e.g. petrol/paraffin/benzene/methylbenzene/ethanol)
  ii. All hydrocarbons (alkanes /alkenes/alkynes)
  iii. Chemicals of life (e.g. proteins, carbohydrates, lipids, starch, sugar)
- **Conductors** these are substances that allow electric current to pass through them.eg. all metals (copper, zinc ,sodium ,magnesium ,aluminum etc) and graphite.
- **Non conductor** are substances that do not allow electric current to pass through them e.g. rubber, wood ,plastic, Sulphur , sand etc.
- **Electricity –**this refers to flow of current
- **Current –**flow of charged particle (electrons or ions)
- **Electrode –**graphite or metal rod dipped in an electrolyte to complete the circuit. An electrode that does not influence/alter the products of electrolysis is called an **inert electrode.** Common inert electrodes include:
  i. **Platinum**
  ii. **Carbon graphite**
- Platinum is not usually used in a school laboratory because it is very **expensive**. Carbon graphite is **easily**/readily and **cheaply** available (from used dry cells).
- ***Cathode –****electrode connected to the negative terminal*
  -electrode over which reduction takes place
- ***Anode –****electrode connected to the positive* terminal
  -electrode over which oxidation takes place.
- **Cation –**positively charged ion
- **Anion –**negatively charged ion

INTRODUCTION

Ionic compounds –conduct heat and electricity because they contain mobile ions, however ionic compounds do not conduct electricity in solid state because the ions are fixed./not mobile

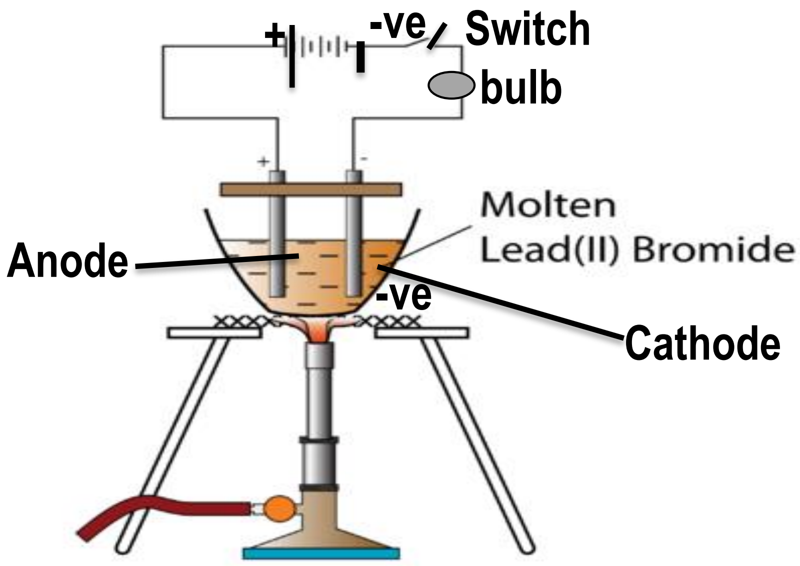
When an electric current is passing through an electrolyte the compound decomposes to constituent ions

The cations move to the anode, gain electrons and become discharged. The anions go to the anode and lose electrons and become discharged

Electrolysis of binary compounds

Binary compound is a compound made of only two ions in molten form

\\\\



ELECTROLYSIS OF MOLTEN LEAD (II) CHLORIDE

When an electric current is passing through molten lead (ii) chloride it decomposes to Pb^{2+} and $2Cl^-$ ions

Reaction at the electrodes

Cathode reaction

Lead ions move to the cathode and gain two electrons and become discharged

$$Pb^{+} + 2e^{-} \rightarrow Pb_{(s)}$$

***Anode* reaction**

Chlorine ions move to the anode lose electrons and become discharged to form chlorine gas

$$2Cl_{(l)}^{-} \longrightarrow Cl_{2(g)} + 2e^{-}$$

State and explain the observation at the cathode

At the cathode

A grey solid is deposited. This is due to discharge of pb^{2+} to lead metal which is grey in colour

***At the* anode**

A green gas is evolved at the anode. This is due to discharge of Cl^- to chlorine gas which is green in colour.

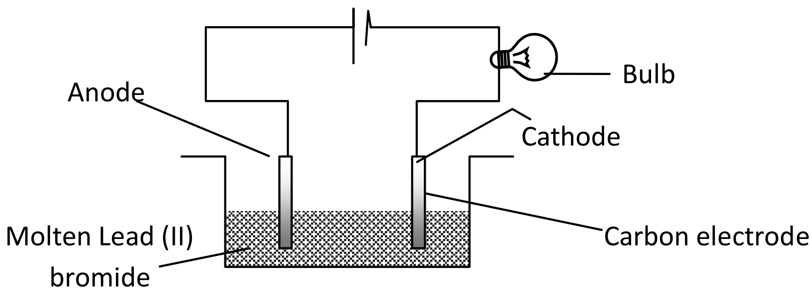
Uses of electrolysis

- Anodizing of Aluminum
- Electroplating -this is coating of a metal with another metal using electric current
- Extraction of very reactive metals like sodium and aluminum from molten ores.
- Purification of metals e.g. copper
- Sacrificial protection

Binary compound	Reaction at cathode	Anode reaction	Observation at the electrodes
Molten NaCl	$Na^{+} + e^{-} \longrightarrow Na_{(s)}$	$2Cl^{-} \rightarrow Cl_{2(g)}$ +2e	Cathode – grey solid deposited Anode –green yellow gas
Molten Al_2O_3	Al^{3+b}+3e ⟶ $Al_{(s)}$	$2O^{2-} \rightarrow O_{2(g)}$ +2e	Cathode – grey solid deposited Anode –colourless gas that relights a glowing splint
Molten $CuCl_2$			
Molten MgO			
Molten $PbBr_2$			
Molten $MgCl_2$			
Molten CuO			
Molten $CaCl_2$			
Molten PbO			

REVISION QUESTIONS ON EFFECT OF AN ELECTRIC CURRENT ON SUBSTANCES

1. The set-up was used to electrolyzed Lead (II) bromide. Study it and answer the questions that follow;



a) Write an ionic equation for the reaction that occurred at the cathode and anode

b) State and explain what happened at the anode and cathode

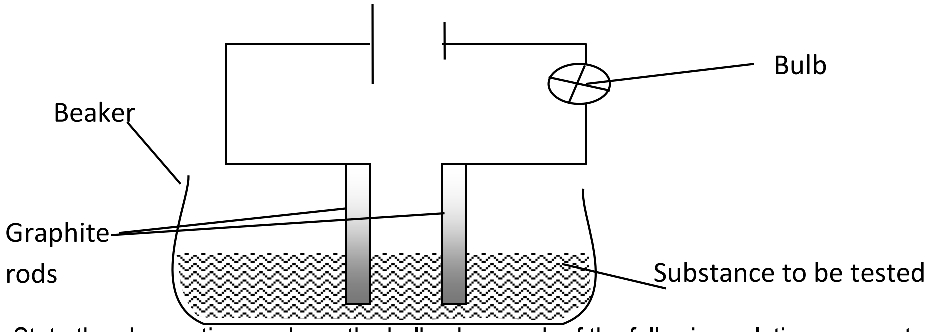
2. When an electric current was passed through two molten substances **E** and **F** in separate voltammeters. The observations recorded below were made: -

Substance	Observation	Type of structure
E	Conducts electric current and a gas is formed at one of the electrodes	
F	Conducts an electric current and is not decomposed	

Complete the table above

3. a). Differentiate the following terms: Electrolyte and non-electrolyte

b). The diagram below is a set-up used to investigate the conductivity of electric current by some aqueous solution. Study it and answer the questions that follow;



a) i). State the observation made on the bulb when each of the following solution were put onto the beaker
Sugar solution

Salt solution

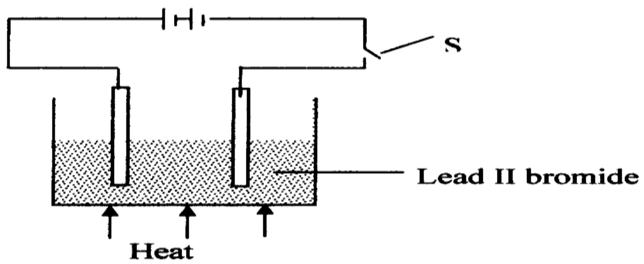
ii). Classify the substance in (a) (i) above as either electrolyte or non-electrolyte

b) If in the above set-up of apparatus, the substance to be tested is Lead (II) Bromide
i). What modification should be included in the set-up?

ii). Write an Ionic equation at the electrodes and state the observation: -
Anode

Cathode

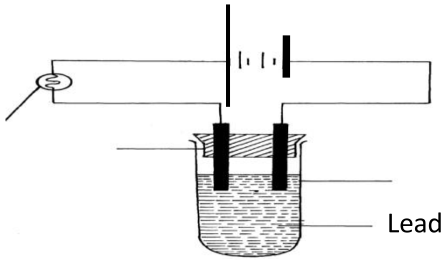
4. The diagram below shows the set up used to investigate the effect of an electric current on molten lead (II) bromide



a) Explain what happens to the lead II bromide during electrolysis

b) Why is it important to carry out the experiment in a fume chamber?

5. The diagram show an experiment for investigating electrical conduction in lead (II) fluoride. Study it and answer the questions that follow:



a) On the diagram
   i. Label the anode and the cathode
   ii. Show the direction of movement of electrons
b) Complete the diagram by indicating the condition that is missing but must be present for electrical conduction to take place.
c) Why is it necessary to leave a gap between the cork and the boiling tube?

d) State the observations that are expected at the electrodes during electrical conduction and at the experiment

e) Write equations for the reactions that take place at the electrodes

f) Why should this experiment be carried out in a fume chamber?

6. The table below shows the electrical conductivity of substance **A, B** and **C**

Substance	Solid state	Molten state	Aqueous solution
A	Conducts	Conducts	Not soluble
B	Doesn't conduct	Conducts	Conducts
C	Doesn't conduct	Doesn't conduct	Not soluble

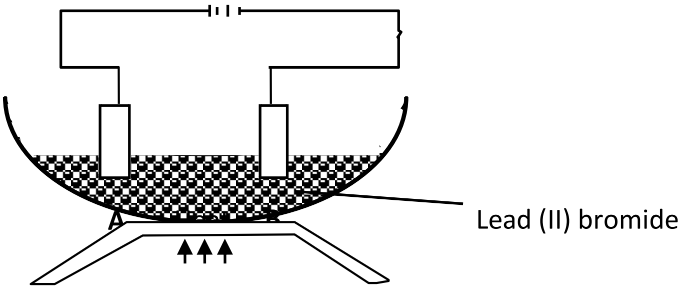
a) Which one of the substances is likely to be plastic?

b) Explain why the substance you have given in **(a)** above behaves in the way it does

c) Which of the substances is likely to be sodium chloride? Explain

d) Give the type of structure and bonding that is present in substance **A**

7. Study the diagram below and use it to answer the questions that follow:-



a) Identify electrodes **A** and **B**

b) Name the product formed at the anode

c) Write the electrode half equation of reaction at electrode **A**

8. Explain the differences in electrical conductivity between melted sodium chloride and liquid mercury

CHAPTER SEVEN: CARBON AND ITS COMPOUNDS

Specific Objectives

a) By the end of this topic, the learner should be able to
b) define allotropy and allotrope
c) explain the physical properties of the carbon allotropes in terms of bonding and how the properties are related to the uses of the allotropes
d) describe some chemical properties of carbon
e) describe laboratory preparation and properties of carbon (IV) Oxide
f) state and explain the physical and chemical properties of carbon(IV) Oxide
g) describe laboratory preparation and some properties of Carbon (II) Oxide
h) describe the chemical reactions of carbonates and hydrogen carbonate
i) describe the manufacture of sodium carbonate
j) explain the advantages and disadvantages of Carbon(IV) oxide and carbon(II) oxide gases in the atmosphere
k) **explain the importance of carbon compounds in the natural environment and industry.**

- Carbon cycle
- Soft drinks manufacture
- Fire extinguishers
- The effects of Carbon(IV) oxide and carbon(II) oxide on the environment

INTRODUCTION

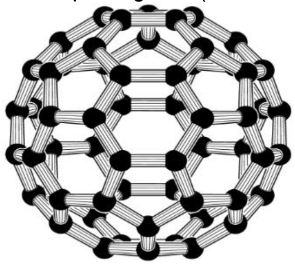
- **Carbon** is a non-metal element with four electrons in its outer most energy level
- It has a valency of 4 and usually forms covalent bonds when combining with other elements ,this is because a lot of energy is involved to gain or lose four electrons hence forms compounds by sharing electrons.
- Has a unique property in that its atoms are bonded together to form long chains. This property is called **catenation**
- *Carbon allotropy –allotropy* is the ability of an element to have **different crystalline forms** but in the same physical state

ALLOTROPES OF CARBON

- *Allotropes –these are different **crystalline*** **forms** of the same element but in same physical state
- There are three main allotropes of carbon
  i. Graphite
  ii. Diamond
  iii. **Fullerene/**Buckminsterfullerene
- Structure, properties and uses of graphite and diamond were tackled under structure and bondin
- So, in this section we are only going to discuss the third allotrope; **Fullerene**

FULLERENE/BUCKMINSTERFULLERENES

A fullerene is an allotrope of carbon in the form of a hollow sphere, ellipsoid, tube, and many other shapes. Spherical fullerenes, also referred to as or buckyballs, resemble the balls used in association football. Cylindrical fullerenes are also called carbon nanotubes(buckytubes). Fullerenes are similar in structure to graphite, which is composed of stacked graphene sheets of linked hexagonal rings. Unless they are cylindrical, they must also contain pentagonal (or sometimes heptagonal) rings



Buckminsterfullerene C_{60}

Note :The discoverers of the Buckminsterfullerene (C_{60}) allotrope of carbon named it after Richard Buckminster Fuller.
As the discovery of the fullerene family came after buckminsterfullerene, the shortened name 'fullerene' is used to refer to the family of fullerenes. The suffix "-ene" indicates that each C atom is covalently bonded to three others (instead of the maximum of four), a situation that classically would correspond to the existence of double covalent bonds

OTHER FORMS OF CARBON AND THEIR USES

- Animal charcoal,wood charcoal, sugar charcoal ,coke ,soot and lamp black are other forms of carbon without definite crystalline structures.they are called Amorphous carbon
- These are form of carbon which do not have distinctive shapes like diamond and graphite.
- These are minute fragments of graphite

Properties of amorphous carbon

- Fairly good conductor of electricity due to presence of graphite
- The charcoal is amorphous light and porous

Uses of amorphous carbon

- Animals charcoal is used to absorb the brown coloring matter in brown sugar which is then turned to white
- Lamp black is used to make shoe polish ,printing ink,paint and reinforcing rubber
- Used as a reducing agent in extraction of metals
- Charcoal are used as fuel for domestic use.

Physical Properties of carbon

i. It occurs widely as a black solid.
ii. Insoluble in water but soluble in organic solvents.
iii. It is poor conductor except the allotrope of graphite.

Chemical Properties of carbon

a) combustion

Carbon burn in air with a red glow to give:

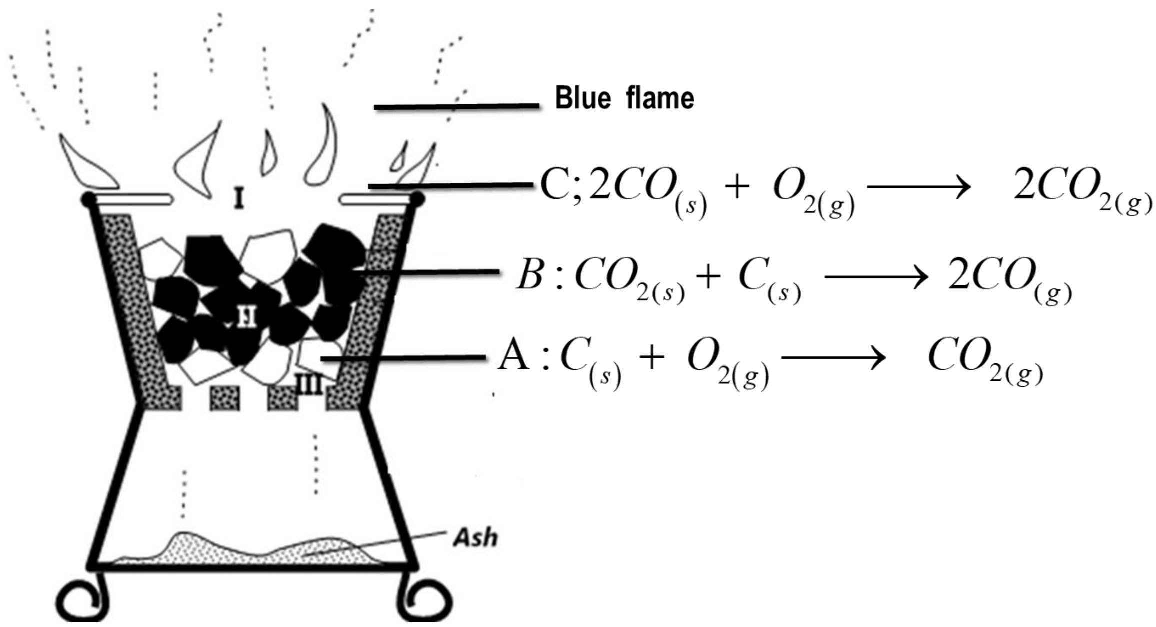
carbon (iv) oxide in plenty of air

$C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2(g)}$

Carbon burns in limited air to form carbon (II) oxide

$2C_{(s)} + O_{2(g)} \longrightarrow 2CO_{(g)}$

carbon (II) oxide and carbon (iv) oxide are also formed during burning of charcoal in a jiko.



In part A there is sufficient supply of air therefore carbon burns completely to form carbon (IV) oxide

$C_{(s)} + O_{2\ (g)} \longrightarrow CO_{2(g)}$

In part B CO_2 is reduced by hot carbon (charcoal) forming carbon (II) oxide

$CO_{2(s)} + C_{(s)} \longrightarrow 2CO_{g)}$

In part C the CO burns in oxygen to form carbon (IV) oxide

$CO_{(s)} + O_{2\ (g)} \longrightarrow 2CO_{2g)}$

NB:A Charcoal stove should not be used in poorly ventilated room since the carbon (IV) Oxide produced will be reduced by hot charcoal to CO which is highly poisonous.

b) Carbon as a reducing agent.

Carbon reduces hot metal oxides to form corresponding metal and carbon (iv) oxide gas

$$2PbO_{(s)} + C_{(s)} \longrightarrow 2\ Pb_{(s)} + CO_{2\ (g)}$$

$$2CuO_{(s)} + C_{(s)} \longrightarrow 2\ Cu_{(s)} + CO_{2\ (g)}$$

$$2ZnO_{(s)} + C_{(s)} \longrightarrow 2\ Zn_{(s)} + CO_{2\ (g)}$$

$$2Fe_2O_{3(s)} + 3C_{(s)} \longrightarrow 4Fe_{(s)} + 3CO_{2\ (g)}$$

Reaction with concentrated Nitric (V) acid and sulphuric (V) acid.

It is oxidized by these acids to form carbon (IV) oxide.

$C_{(s)} + 2H_2SO_{4\ (l)} \longrightarrow CO_{2(g)} + 2SO_{2(g)} + 2H_2O_{(l)}$

Observations

- ✓ *black solid dissolves*
- ✓ *A gas with a pungent smell is produced*

$C + 4HNO_3 \longrightarrow CO_2 + 4NO_2 + 2H_2O$

Observations

- ✓ *brown fumes of nitrogen (IV) oxide produced*
- ✓ *Black solid dissolves*

OXIDES OF CARBON

There are two oxides of carbon;

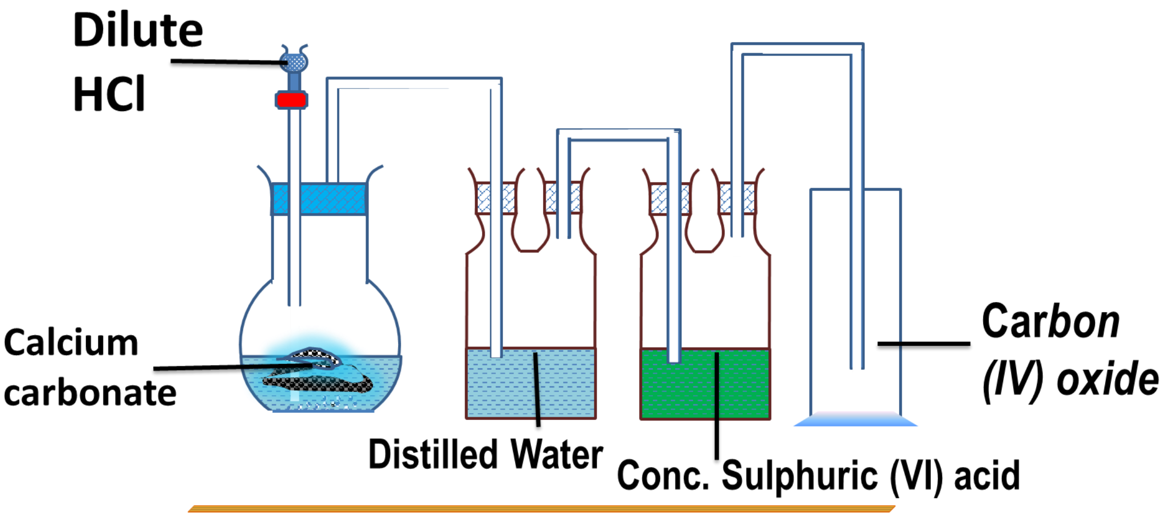
Carbon (II) oxide and Carbon (IV) oxide

CARBON (IV) OXIDE

Laboratory preparation

carbon (IV) oxide is prepared in the laboratory by reacting a carbonate with a suitable acid, preferably Calcium carbonate and dilute hydrochloric acid

laboratory preparation of preparation carbon (iv) oxide



Sodium hydrogen carbonate or water absorbs traces of HCl fumes and concentrated H_2SO_4 dries the gas
The gas is collected by downward delivery as it is denser than air.

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2\,(S)} + CO_{2(g)} + H_2O_{(l)}$$

Dried by conc. H_2SO_4 or $CaCl_2$

Note .*$CaCO_3$ and dilute H_2SO_4 cannot be used to prepare CO_2 due to formation of* ***insoluble $CaSO_4$*** *which* ***coats the carbonate*** *preventing further reaction between the acid and carbonate. For the same reason $PbCO_3$ and dilute HCl, $BaCO_3$/$PbCO_3$ and dilute H_2SO_4 cannot be used. $CaCO_3$ is preferably used because it is cheap and easily available*

Physical Properties of carbon (IV) oxide

- Colourless and odourless
- Slightly soluble in water
- Denser than air hence collected by downward delivery

CHEMICAL PROPERTIES OF CARBON (IV) OXIDE

a) Extinguishes a burning splint because it does not burn and does not support combustion

b) Reaction with water

Reacts with water forming weak carbonic acid

$$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$$

Carbon (IV) oxide changes moist blue litmus paper red.

c) Reaction with *lime water (calcium hydroxide)*

If carbon (IV) is bubbled through Lime water it forms a white precipitate. If more CO_2 is bubbled through the white precipitate , the white precipitate dissolves to form a colourless solution, the white precipitate reappears on heating. This is confirmatory test for carbon (IV) oxide

Explanation: when carbon (IV) oxide is bubbled through lime water, it reacts to form insoluble $CaCO_3$ (white precipitate) which reacts with more of carbon (IV) oxide to form Ca $(HCO_3)_2$ on heating the Ca $(HCO_3)_2$ decomposes to give out insoluble $CaCO_3$ as a white precipitate.

$$Ca\ (OH)_{2(aq)} + CO_{2(g)} \longrightarrow CaCO_{3(s)} + H_2O_{(l)}$$

$$CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \longrightarrow Ca(HCO_3)_{2(aq)}$$

$$Ca\ (HCO_3)_{2(s)} \xrightarrow{heat} CaCO_{3(s)} + CO_{2\ (g)} + H_2O_{(l)}$$

Raction with alkalis

CO_2 reacts with alkalis forming corresponding carbonates, but if excess CO_2 is bubbled through alkali metal hydrogen carbonate is formed

$$2NaOH\ + CO_2 \longrightarrow Na_2CO_{3(s)} + H_2O_{(l)}$$

$$Na_2CO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \longrightarrow 2NaHCO_{3(aq)}$$

)

$$2KOH_{(aq)} + CO_{2(g)} \longrightarrow K_2CO_{3(s)} + H_2O_{(l)}$$

$$K_2CO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \longrightarrow 2KHCO_{3(aq}$$

d)* Reaction with burning *elements

When burning magnesium is lowered into a gas jar full of* CO_2 *it continues to burn. Forming white solid and black specs of carbon. This is because the heat produced by a burning magnesium is so high that it decomposes CO_2 into carbon and oxygen gas which oxidizes the metal to white magnesium oxide and carbon (IV) oxide is reduced to carbon (black solid

Observations: white solid and black specs of carbon are formed

$$CO_{2(g)} \xrightarrow{heat\ from\ hot\ Mg} C_{(s)} + O_{2(g)}$$

$$Mg_{(s)} + O_{2(g)} \longrightarrow MgO_{(s)}$$

Overall equation

$$2Mg_{(s)} + CO_{2(g)} \longrightarrow 2MgO_{(s)} + C_{(s)}$$

Uses of carbon (IV) oxide

- In manufacture of baking powder. Baking powder contain sodium hydrogen carbonate and a solid acid such as tartaric acid. these compounds react together in presence of water to form carbon (IV) oxide. further gas is produced during cooking due to decomposition of $NaHCO_3$.in presence of heat in bread making yeast causes carbon (IV) oxide to be produced causing the dough to swell
- Carbonate's derivatives
- Manufacture of sodium carbonate in **Solvay process**
- Used in soft drinks e.g soda . in these drinks the gas is dissolved under pressure.
- Used in Fire extinguishers since it is denser than air and does not support combustion.

- As a refrigerant (as in dry ice because it is very cold and sublimes leaving no residue and does not

Roles of CO_2 in soft drinks;
- Acts as a preservative
- It gives taste/adds flavor

When a bottle containing soft drink is opened there is effervescence/ bubbles of a gas is observed as CO_2 dissolved under pressure is released
What is the role of tartaric acid in baking powder?
Tartaric acid reacts with $NaHCO_3$ in the baking powder in presence of water to produce carbon (IV) oxide causing the dough to rise

cause wetness)Making rain during drought in little rain drop

When cooking mandazi using baking powder the dough rises when put in a frying pan explain.

Baking powder contains **$NaHCO_3$**, so when the dough is put in a frying pan the $NaHCO_3$ **decomposes on heating to produce carbon (IV) oxide** and it is the **CO_2 which makes the dough** to rise **as it forces its way out**. That is why mandazi always have cracks.

What is the role of tartaric acid in baking powder?

Tartaric acid reacts with $NaHCO_3$ in the baking powder **in presence of water** producing carbon (IV) oxide which makes the dough to rise.

State the properties of carbon (IV) oxide that make it suitable to be used in fire extinguishers

- it is denser than air
- it does not burn
- it doesn't support burning
- it is not poisonous

Fire extinguishers

Carbon (IV) oxide is used to put out fires especially those caused by petrol. Oil and electricity. If water is used to put out oil fires, the burning oil will float on the water. This means that the fire would continue to burn on the surface. Carbon (IV) oxide is used to put out electric fires since water conducts electricity.

Types of fire extinguishers

- **Soda acid extinguisher-** contains sodium hydrogen carbonate at the bottom and dilute sulphuric (VI) acid on top. The two react to produce carbon (IV) oxide
- **Foam extinguisher**. Contains $NaHCO_3$ solution and Aluminium sulphate which is acidic when these two react they form a foam of CO_2.
- **Carbon (IV) oxide extinguisher** contains liquid carbon (IV) oxide under pressure.

CARBON (II) OXIDE

Laboratory preparation of Carbon (II) oxide

1. ***Preparation from methanoic acid***

Action of concentrated sulphuric acid on methanoic acid/ethanedioc acid

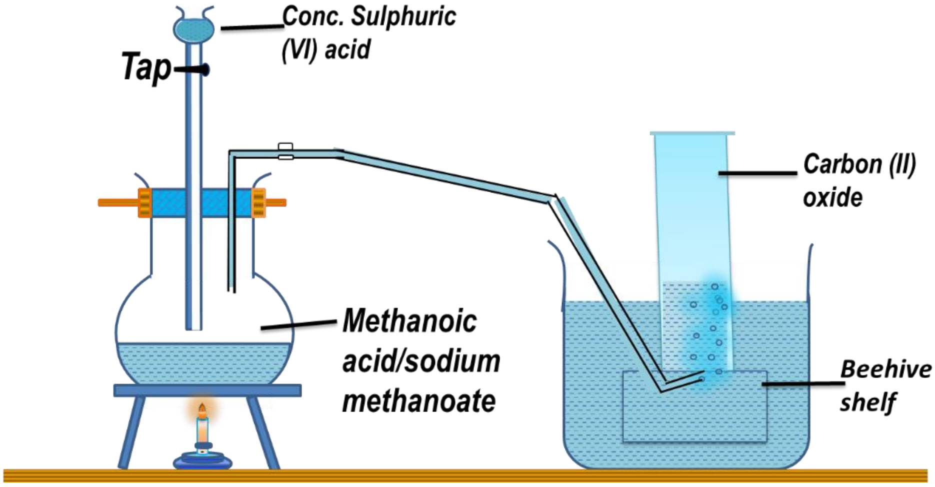
$$HCOOH_{(l)} \xrightarrow{conc.H_2SO_4} CO_{(g)} + H_2O_{(l)}$$

Sodium methanolate can also replace methanoic acid

$$HCOONa_{(s)} + H_2SO_{4(l)} \longrightarrow NaHSO_{4(aq)} + CO_{(g)} + H_2O_{(l)}$$

NB: concentrated Sulphuric (VI) acid acts as a dehydrating agent.

Preparation of carbon (II) oxide from dehydration of methanoic acid

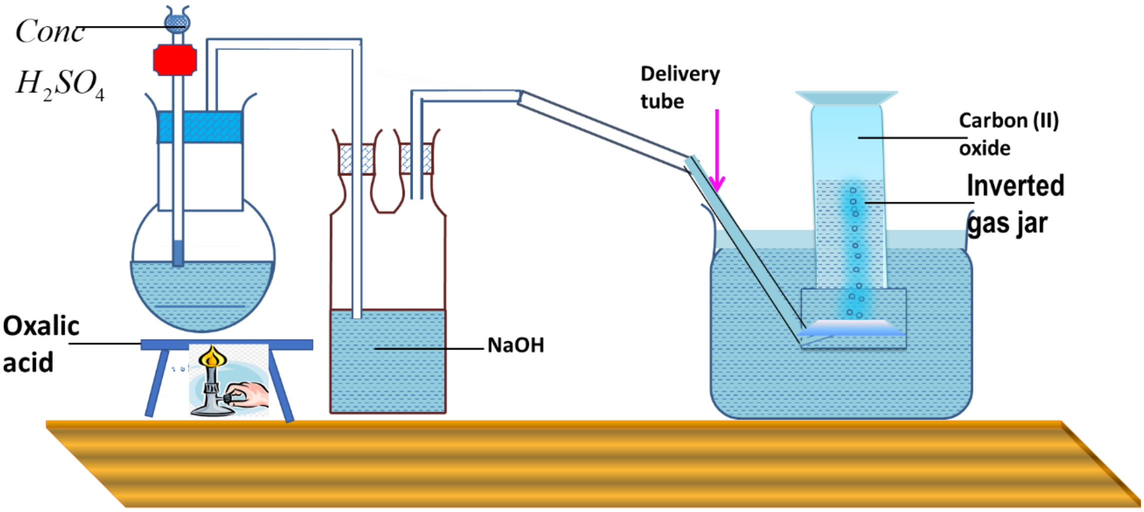


2. **Preparation from Ethanedioc acid (oxalic acid)**

When oxalic acid is used both carbon (IV) oxide and carbon (II) oxide gases are produced and the gases are passed through a concentrate solution of a strong alkali NaOH, KOH or $Ca(OH)_2$ to absorb co_2

$$H_2C_2O_{4_{(s)}} \xrightarrow{conc.\ H_2SO_4} CO_{2(g)} + CO_{(g)} + H_2O_{(l)}$$

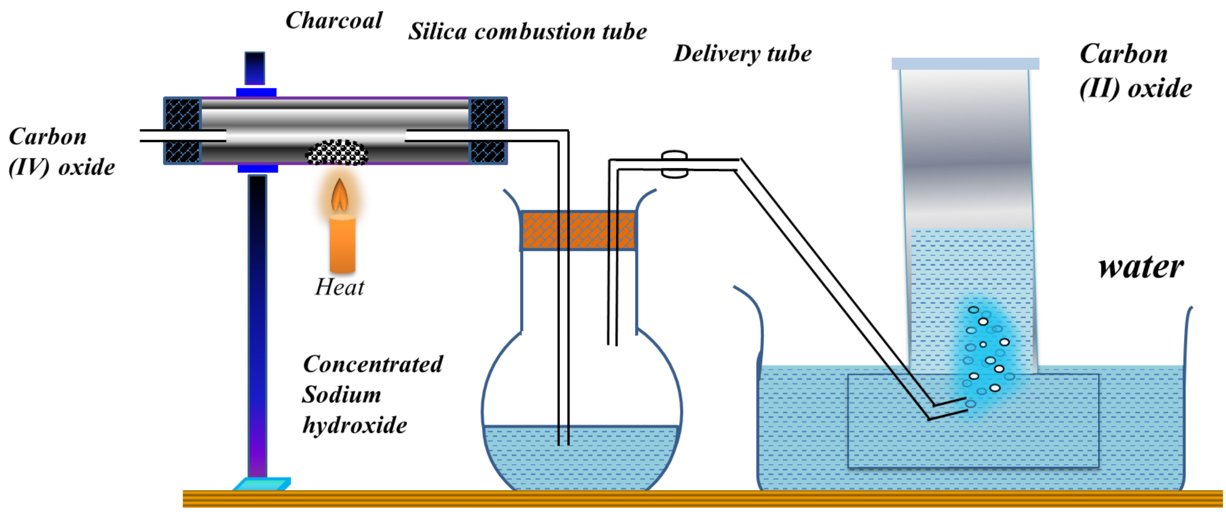
Preparation of carbon (II) oxide from dehydration of oxalic acid



3. <u>Preparation from carbon (IV) oxide</u>

Carbon (II) oxide is prepared from reduction of carbon (IV) oxide by carbon:the excess carbon (IV) oxide is removed by passing the gases through concentrated sodium hydroxide/potassium hydroxide
hydroxide

Preparation of carbon (II) oxide from carbon (IV) oxide gas and charcoal



Properties of carbon (II) oxide

- Colourless, odourless gas
- Density is slightly lower than air
- It is insoluble in water so it is collected over water
- Tasteless

NB: CO is is known as silent killer because it is odourless and colourless.

Chemical properties of carbon (II) oxide

a) **Neutral to litmus papers**

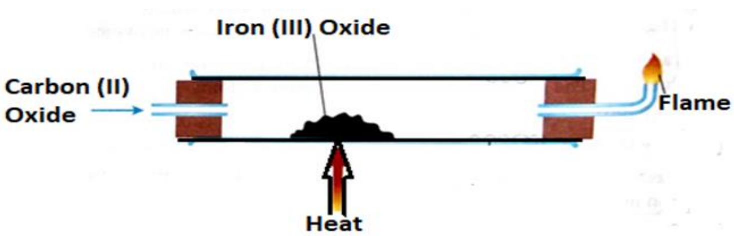
b) **Burning in air**

carbon (II) oxide burns in air with a **blue flame** forming carbon (IV) oxide

$$2CO + O_{2(g)} \rightarrow 2CO_{2(g)}$$

c) **carbon (II) oxide as a reducing agent**

Reduces metal oxides of less reactive elements to free metals



$$PbO_{(s)} + CO_{(g)} \longrightarrow Pb_{(s)} + CO_{2\,(g)}$$

$$Fe_2O_{3(s)} + CO_{(g)} \longrightarrow Fe_2O_{3(s)} + CO_{2(g)}$$

Reacts with iron metal to form carboxyls

$Fe_{(s)} + 5CO_{(g)} \longrightarrow Fe(CO)_5$ (pentacarboxyl iron)

The carbon (II) oxide is highly poisonous.

It combines with haemoglobin, unstable compound is formed which prevents oxygen from being circulated in the body hence suffocation and eventually death result

Uses of carbon (II) oxide

- Widely used in manufacture of alcohol
- Extraction of less reactive metals from their ores
- Used as a fuel

Differences between carbon (IV) oxide and carbon (II) oxide

CO_2	**CO**
Forms white precipitate with lime water	No reaction with lime water
No reaction with metallic oxides	Reduces metal oxide to metals
Not poisonous	Very poisonous
Soluble in water and alkali	In soluble in water and alkali
Acidic	Neutral

LARGE SCALE MANUFACTURE OF SODIUM CARBONATES (NA_2CO_3)(SODA ASH)

- It is obtained in large scale in two methods
- From trona
- Through Solvay process

EXTRACTION FROM TRONA

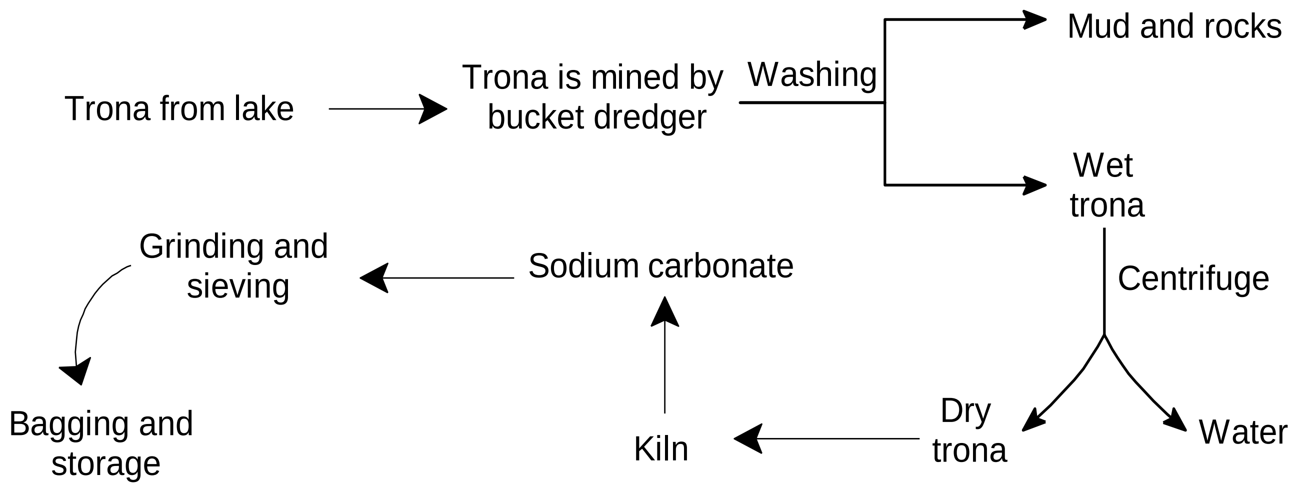
Is a double salt with formula $Na_2CO_3.NaHCO_3.2H_2O$. Trona is extracted using bucket dredgers and then crushed. The crushed trona is washed with water to remove mud and small rocks. It is then centrifuged to remove water of crystallization. It is then roasted in a kiln at 300°C to decompose to Na_2CO_3 as follows. This process is called **thermal decomposition.**

At centrifuge $NaHCO_3.Na_2CO_3.2H_2O_{(s)} \longrightarrow Na_2CO_3.NaHCO_{3(s)} + H_2O_{(l)}$

$$Na_2CO_3.NaHCO_{3(s)} \longrightarrow Na_2CO_{3(s)} + H_2O_{(l)} + CO_{2(g)}$$

Na_2CO_3 is then grinded and bagged for storage and transportation

SCHEMATIC DIAGRAM FOR EXTRACTION OF SODIUM CARBONATE FROM TRONA



After removal of trona from the lake, a solution which is rich in sodium chloride remains. the solution is then pumped into shallow basins where evaporation takes place until the percentage of sodium chloride is 14% , then the solution is transferred to another basin and cooled to crystalize .
The solution which contains both trona and sodium chloride is separated by fractional crystallization.
During the day, when the temperature is about 40^0C and trona crystallizes and is removed . The solubility of Sodium Chloride is high at high temperatures therefore during the day it dissolves while at night when the temperature is about 20^0C sodium chloride crystallizes and is removed.

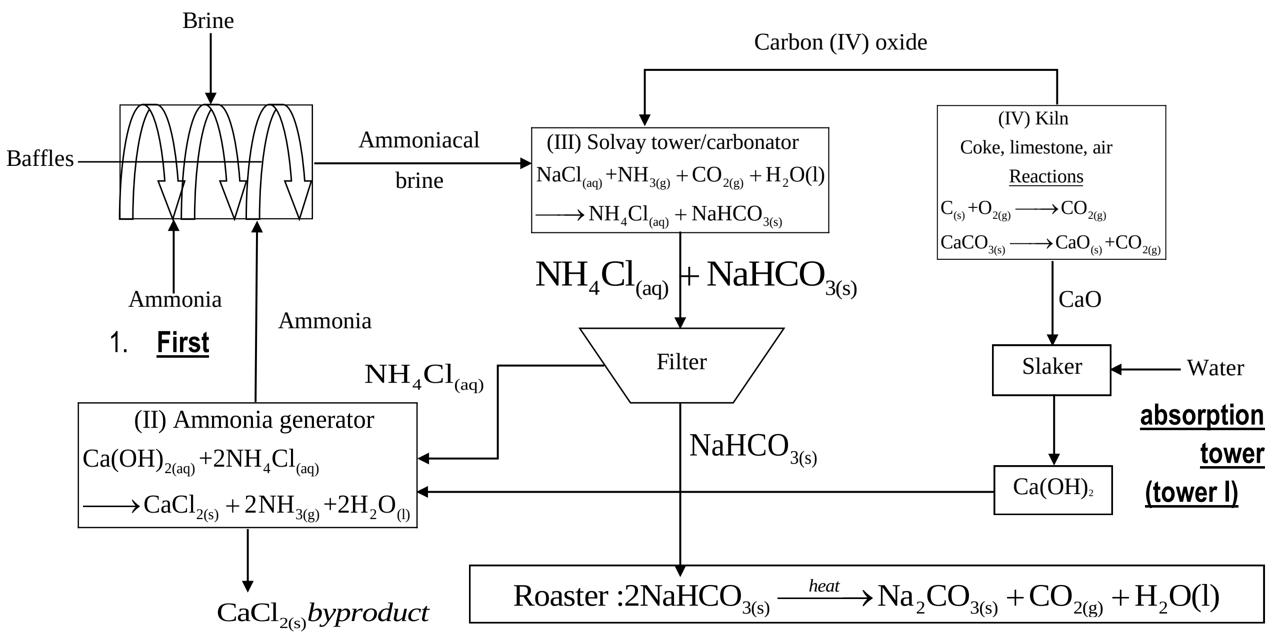
Effects of heat on carbonates and hydrogen carbonates refer to salts

SOLVAY PROCESS

Raw materials needed

- Brine
- $CaCO_3$
- Ammonia
- Coke

SCHEMATIC DIAGRAM FOR MANUFACTURE OF TRONA THROUGH SOLVAY PROCESS



Ammonia is mixed with brine to form ammoniacal brine. The absorption tower is fitted with baffles. which allow the liquids trickle down slowly to allow more time for proper mixing with gases. The baffles also increase the surface area for gases to dissolve in the liquids.

2. **Ammonia Generator**

Ammonia is generated by reacting ammonium Chloride from Solvay tower with calcium hydroxide from the slaker.

$$\mathbf{2NH_4Cl_{(aq)} + Ca(OH)_{2(aq)} \longrightarrow CaCl_{2(aq)} + 2NH_{3(g)} + H_2O_{(l)}}$$

$CaCl_2$ is the only a byproduct in this process and is widely used as a drying agent.

3. **Carbonator/Second Absorption Tower**

Carbonator is where the main reaction takes place, Reaction takes place in two stages as follows

$NH_{3(g)} + H_2O_{(l)} + CO_{2(g)} \longrightarrow NH_4HCO_{3(aq)}$ (this reaction occurs in upper part of Solvay tower)

$NH_4CO_{3\ (aq)} + NaCl_{(aq)} \longrightarrow NaHCO_{3(s)} + NH_4Cl_{(aq}$ (this reaction occurs in lower part of Solvay tower)

Overall equation

$$NaCl_{(aq)} + NH_{3(g)} + CO_{(g)} + H_2O_{(\ l)} \longrightarrow NaHCO_{3(s)} + NH_4Cl_{(aq)}$$

NH_4Cl is more soluble than $NaHCO_3$. These salts are separated through filtration where $NaHCO_3$ is obtained as a residue while NH_4Cl is obtained as a filtrate. NH_4Cl is taken to ammonia generator. Fractional crystallization can also be used to separate the salts because $NaHCO_3$ is less soluble under low temperatures.

$NaHCO_3$ is roasted to obtain Na_2CO_3

$$NaHCO_{3(S)} \xrightarrow{heat} Na_2CO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}$$

NOTE: The reaction in the carbonator is highly exothermic hence water is made to circulate around the carbonator to cool the products. This ensures that $NaHCO_3$ crystallizes out since it is less soluble at low temperature.

4. **Chamber four calcium carbonate kiln**

Carbon (IV) oxide is generated in two ways

$$CaCO_{3(s)} \xrightarrow{heat} CaO_{(S)} + CO_{2(g)}$$

$$C_{(s)}\ (coke) + O_{2(\ g)} \longrightarrow CO_{2(g)}$$

Solvay process is regarded as one of the most economical process because most of the byproducts like ammonia and carbon (IV) oxide are recycled. The raw materials are cheap and the only by-product is calcium chloride which is widely used a drying agent for laboratory gases.

CaO is slaked with water to form $Ca(OH)_2$ which is directed to ammonia generator to form ammonia

NB: problems student face

Most students do not know what reacts at ammonia generator

Most students confuse the second chamber (carbonator) with the fourth chamber (limestone kiln)

Predicted questions

- **Name two uses of $CaCl_2$**
- ✓ fused **$CaCl_2$** Used as a drying agent for gases
- ✓ Used in extraction of sodium metal to lower the boiling point
- ***Name two uses of Na_2CO_3***
- ✓ *Used in manufacture of glass*
- ✓ *Manufacture of detergents*
- ✓ *Used in paper industry*
- ✓ *Used in softening hard water*
- **Name two uses of $NaHCO_3$**

- ✓ Use in manufacture of baking soda
- ✓ Used in manufacture of aerated drinks
- **Give reasons why the Solvay plant should be located near a river**
- ✓ Water is a raw material
- ✓ Water is needed as a coolant as the reaction in the Solvay tower is highly exothermic
- **Name two materials recycled in the Solvay process**
- ✓ Ammonia and carbon (IV) oxide
- **Give the advantage of recycling**
- ✓ It minimizes production/running cost
- ✓ It minimizes pollution

- **State two costs incurred in running the plant**
- ✓ Power cost
- ✓ Maintenance cost
- **Explain why K_2CO_3 is not preferably prepared using this method.**

This is because the $KHCO_3$ is highly soluble at lower temperature hence will not precipitate/be difficult to separate from NH_4Cl . $KHCO_3$ has similar solubility to NH_4Cl at the same temperatures therefore difficult to separate.

- **Name two methods used to separate the products at the Solvay tower**

Fractional crystallization
Filtration

- **Explain why Solvay process is considered as a very economical process**

Raw materials are cheap
Most of the byproducts like CO_2 and ammonia are recycled

POLLUTION EFFECT BY CARBON AND CARBON COMPOUNDS

a) Carbon (IV) oxide from burning fuels accumulates in the upper atmosphere and forms a layer that traps heat energy reflected from the earth raising air temperatures. **This is called global warming**
Global warming leads to melting of glaciers and ice caps leading to climatic changes

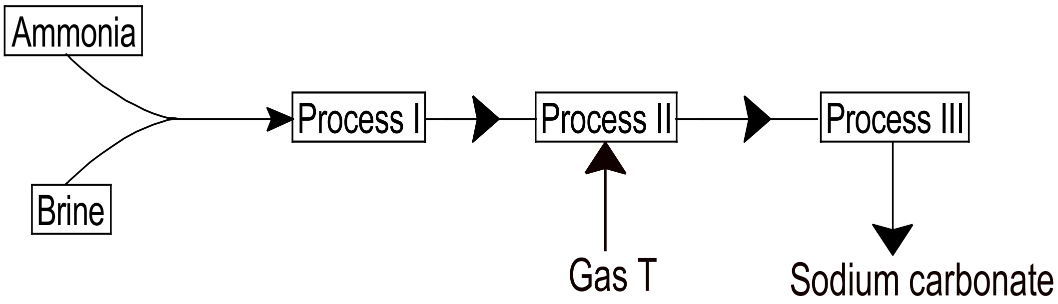
b) Carbon (II) oxide is very poisonous gas and lead to suffocation leading to death of animals by combining with haemoglobin to form a very stable compound called caboxyhaemoglobin which does not dissociate

c) to release oxygen hence leads to suffocation

Note: Green house effect as a pollution effect of CO_2, **is a very weak point** and does not **SCORE (for examination)** but increased green house gases leads to **global warming** which is the scoring point. The effect of CO_2 to bring about acid rain may not also score as carbonic acid is a very weak acid

REVISION QUESTION ON CARBON AND ITS COMPOUNDS

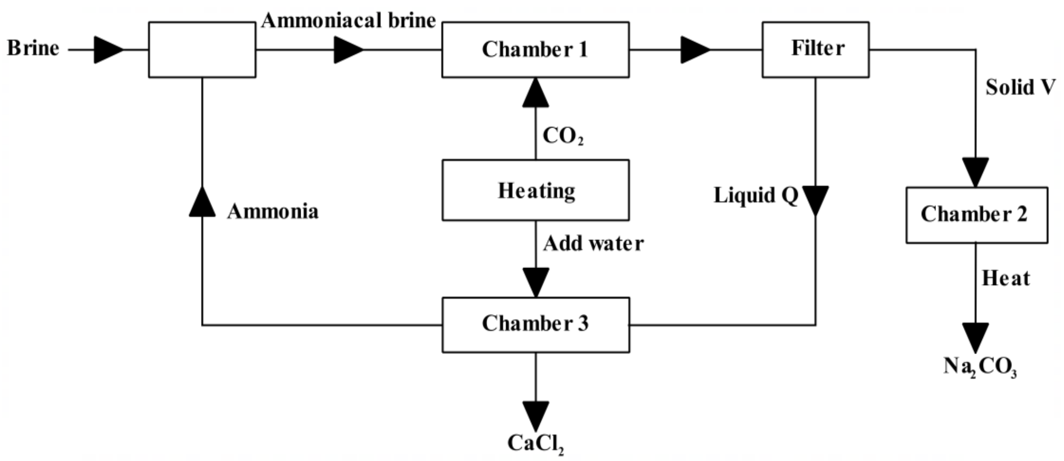
1. Below is a simplified scheme of Solvay process. Study it and answer the questions that follow



a) Identify gas T

b) Write equation for the reaction that takes place in process II and III

2. The diagram below shows the stages in the manufacture of sodium carbonate .study the diagram below and use it to answer the questions that follow



a) Name three starting materials in the manufacture of sodium carbonate

b) Which substances are recycled in this process

c) identify the chambers in which the recycled substances are renegerated

d) Name the substances V and Q

e) Give an equation for the reaction which occurs:

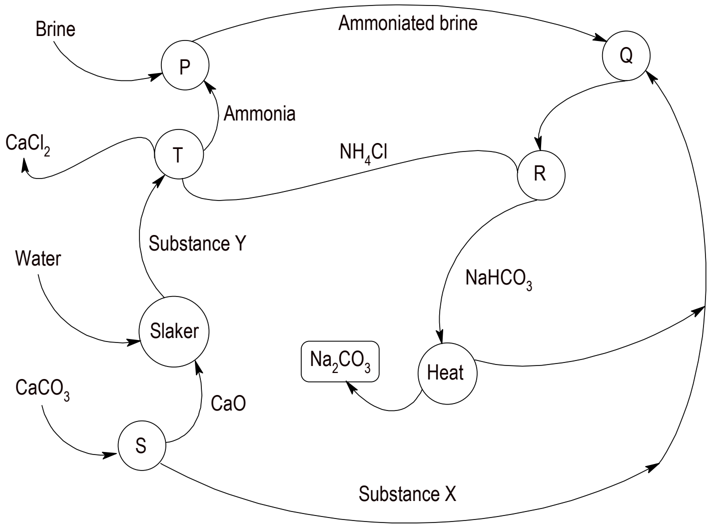
i. In the reaction chamber I

ii. When solid V is heated

iii. In the reaction chamber 3

f) State one commercial use for sodium carbonate

3. Use the flow chart below to answer the questions that follow



(a) Name the substances labelled:(2 marks)

X……………………………………………………………………………………

Y……………………………………………………………………………………

(b) Name 2 substances being recycled in the process represented by the flow chart. (2 marks)

………………………………………………………………………………………

………………………………………………………………………………………

(c) Name the process that takes place in: (2 marks)

S………………………………………………………………………………………

R………………………………………………………………………………………………

(d) Give 2 uses of calcium chloride. (1 mark)

………………………………………………………………………………………………

………………………………………………………………………………………………

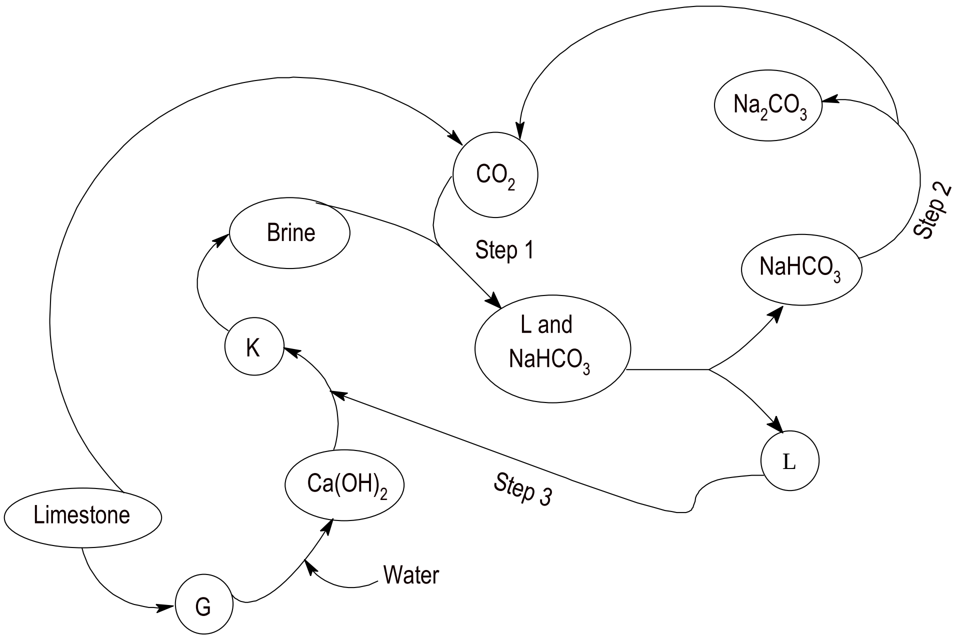
………………………………………………………………………………………………

(e) Write equations for the reaction that take place in (2 marks)
Q

T

(f) Other than softening of hard water give 2 other uses of sodium carbonate (1 mark)

4. The flow chart below shows the stages in the industrial manufacture of sodium carbonate. Study it and answer the questions that follow.



a) Name
K………………………………. ………………………………………………..
L……………………………………………………………………………………..
G………………………………… …………….. (3 marks)

b) What is the name given to industrial manufacture of sodium carbonate (1 mark)

c) Write the equations of reaction taking place in **step I** (2 marks)

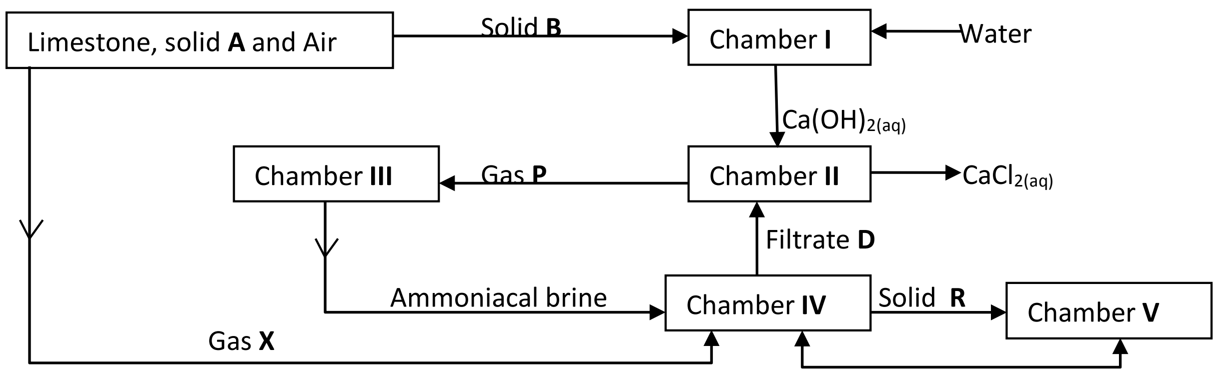
d) How is **step 2** achieved (1 mark)

e) State how the products in **step 1** are separated (2 marks)

f) This process is said to be very efficient. Explain this(2 marks)

g) Using dots and crosses show how the bonding of carbon (IV) oxide is achieved (atomic numbers C- 6, O-8) (2 marks)

5. The flow chart below shows how sodium carbonate is manufactured by the solvary process. Study it and answer the questions that follow;



a) Name the following substances; (3 marks)
   i. Gas **P** ……………………………………………………………………………
   ii. Gas **X** ……………………………………………………………………………
   iii. Solid **A** ……………………………………………………………………………
   iv. Solid **B** ……………………………………………………………………………
   v. Filtrate **D** ……………………………………………………………………………
   vi. Solid **R** ……………………………………………………………………………

b) Write chemical equations for the reactions that occur in;
   chamber **I** (1 mark)

   Chamber **II** (1 mark)

c) Name **two** gases that are recycled in the above process (2 marks)

d) Give **one** use of sodium carbonate (1 mark)

6. Both graphite and molten lead (II) chloride conduct electricity. State how each of the substances conducts electricity.
   a) Graphite (1 mark)

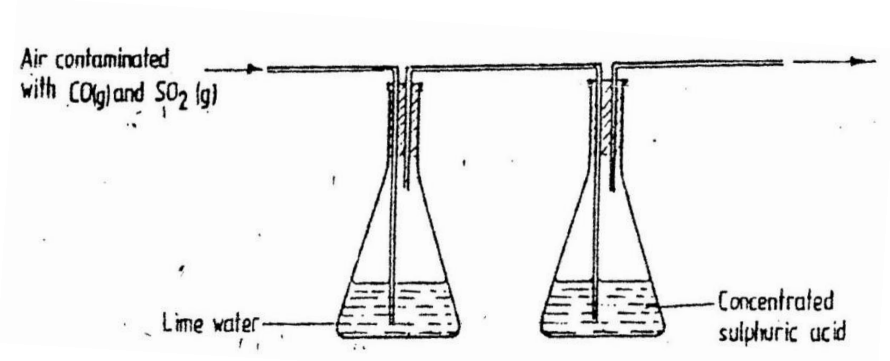
   b) Molten lead (II) chloride (1 mark)

7. When a salt T is heated, a black solid is left and a colourless gas which forms a white precipitate with calcium hydroxide solution is evolved. Identify T and write an equation for the decomposition. (2 marks)

8. What is meant by the term allotropy? Give an example of an element that exhibits allotropy. (1mk)

9. Give a reason why calcium hydroxide solution is used to detect the presence of carbon (IV) oxide gas while sodium hydroxide solution is NOT (1 mark)

10. A sample of air contaminated with carbon (II) oxide and sulphur (VI) oxide was passed through the apparatus shown in the diagram below.



Which contaminant was removed by passing the contaminated air through the Apparatus. Explain (2 marks)

11. Explain how you would obtain solid sodium carbonate from a mixture of lead carbonate and sodium carbonate powders. (3 marks)

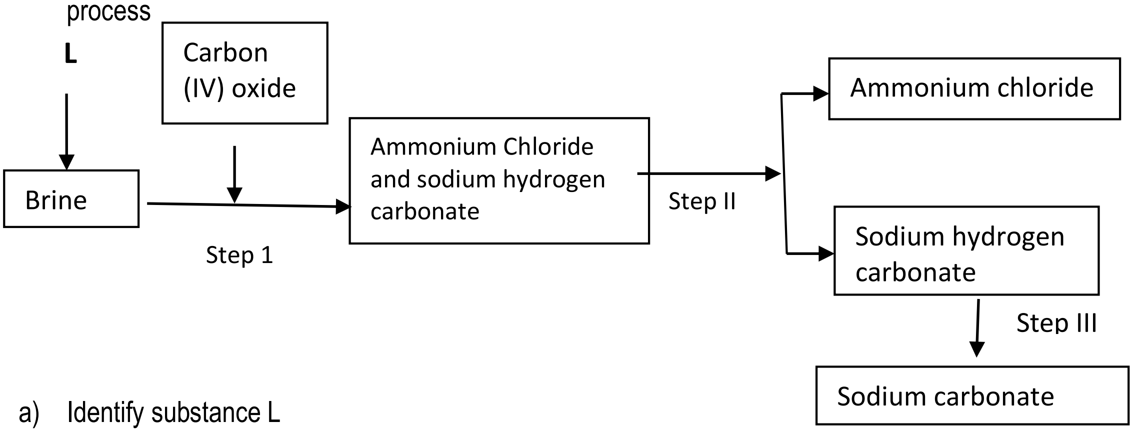
12. When extinguishing a fire caused by burning kerosene, carbon (IV) oxide is used in preference to water .Explain

13. When excess carbon (II) oxide gas was passed over heated lead (II) oxide in combustion tube, lead (II) oxide was reduced
    a) Write an equation for the reaction, which took place

    b) What observation was made in the combustion tube when the reaction was complete?

    c) Name another gas, which could be used to reduce lead (II) oxide

14. The simplified flow chart shows some of the steps in the manufacture of sodium carbonate by the Solvay process

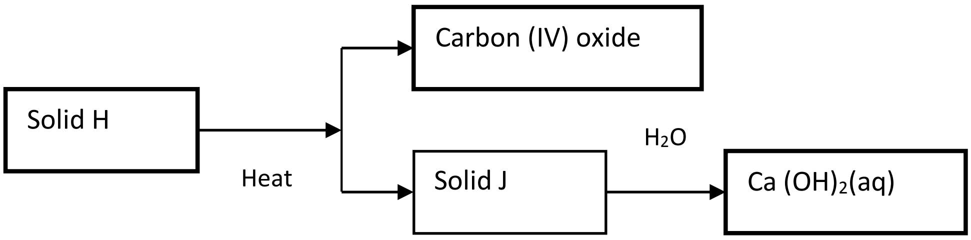


a) Identify substance L

b) Name the process – taking place in step II

c) Write an equation for the reaction, which takes place in step III

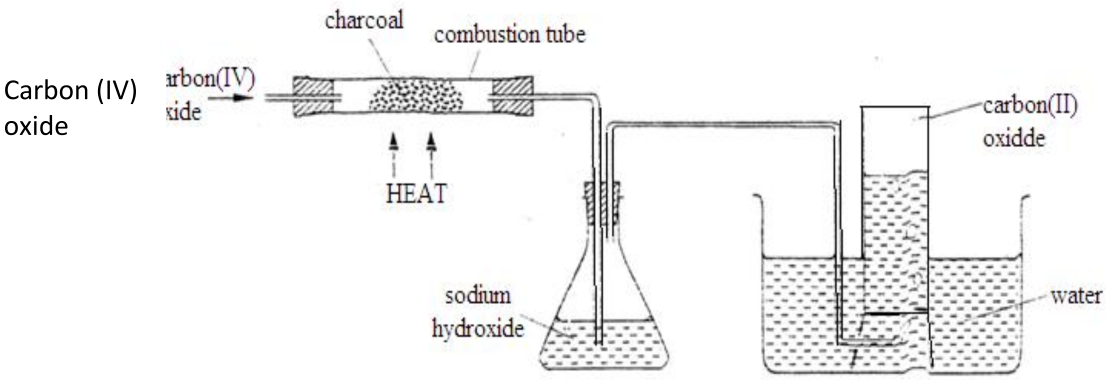
15. Use the scheme below to answer the questions that follow



a) Identify the solid
H ....
J ....

b) State one commercial use of solid J

16. In an experiment, carbon (IV) oxide gas as passed over heated charcoal and the gas produced collected as shown in the diagram below



a) Write an equation for the reaction that took place in the combustion tube

b) Name another substance that can be used instead of sodium hydroxide

c) Describe a sample chemical test that can be used to distinguish between carbon (IV) oxide and carbon (II) oxide

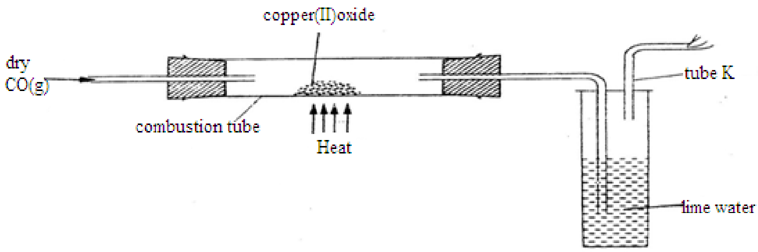
d) Give one use of carbon (II) oxide

17. When the oxide of element H was heated with powdered carbon the mixture glowed and carbon (IV) oxide was formed. When the experiment was repeated using the oxide of element J, there was no apparent reaction.

a) Suggest one method that can be used to extract element J from its oxide

b) Arrange the elements H, J and carbon in the order of their decreasing creativity.

18. The apparatus shown below shown below was used to investigate the effect of carbon (II) oxide on copper (II) oxide.



a) State the observation that was made in the combustion tube at the end of the experiment.

b) Write an equation for the reaction that took place in the combustion tube

c) Why is it necessary to burn the gas coming out of tube K?

19. When carbon (IV) oxide gas was passed through aqueous calcium hydroxide a white precipitate was formed

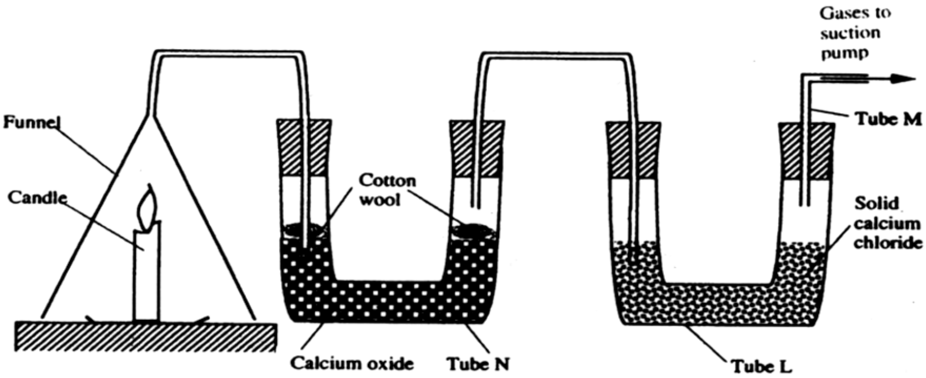
a) Write an equation for the reaction that took place

b) State and explain the changes that would occur when carbon (IV) oxide gas is bubbled through the white suspension

20. a). Candle wax is mainly a compound consisting of two elements. Name the two elements (2 marks)

b). The set- up below was used to investigate the burning of a candle study it and

answers the questions that follow



i. What would happen to the burning candle if the pump was turned off? Give reasons

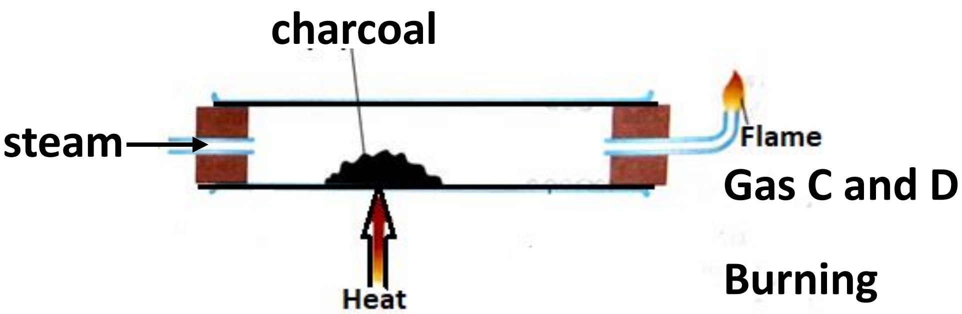
ii. State and explain the changes in mass that are likely to occur in tube N by the end of the experiment (3 marks)

iii. Name two gases that come out through tube M (2 marks)

iv. Name another substance that could be used in the place of calcium oxide in tube N

21. Give the role of carbon (IV) oxide in carbonated drinks (2 marks)

22. When steam was passed over heated charcoal as shown in the diagram below, hydrogen and carbon (II) oxide gases were formed



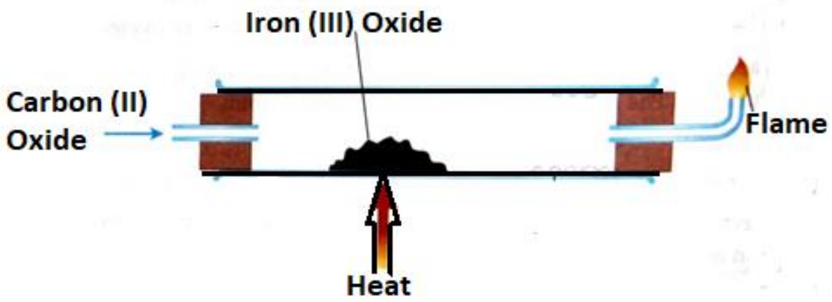
a) Identify Gas C and D

b) Write the equation for the reaction which takes place (1 mark)

c) Name two uses of carbon (II) oxide gas, which are also uses of hydrogen gas (2 marks)

23. When a candle was brunt completely. The total mass product was found to be greater than the original mass of the candle. Explain

24. Carbon (II) oxide gas passed over heated Iron (III) oxide as shown in the diagram below.



a) Give the observation made in tube P (1 mark)

b) Write the equation for the reaction which takes place in tube P. (1 mark)

.

25. Dry carbon (II) oxide gas reacts with heated lead (II) oxide as shown in the equation below

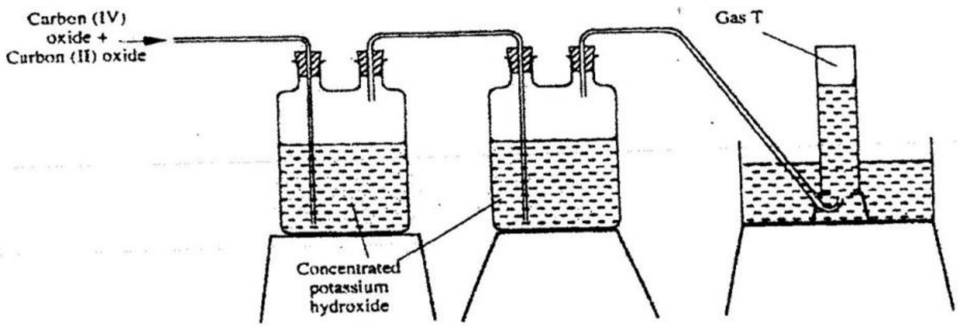
$PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$

a) Name the process undergone by the lead (II) oxide (1 mark)

b) Give a reason for your answer in (a) above (1 mark)

c) Name another gas that can be used to perform the same function as carbon (II) oxide gas in the above reaction.

26. The diagram below represents part of a set – up used to prepare and collect gas T.

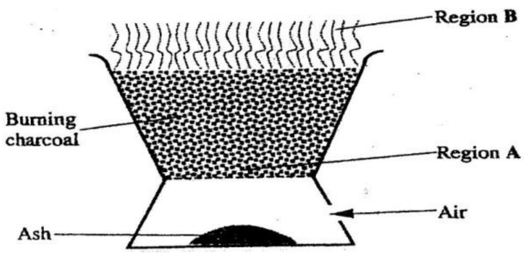


a) Name two reagents that are reacted to produce both carbon (IV) oxide and carbon (II) oxide. (1 mark)

b) Write the equation for the reaction which takes place in the wash bottles. (1 mark)

c) Give a reason why carbon (II) oxide is not easily detected. (1 mark)

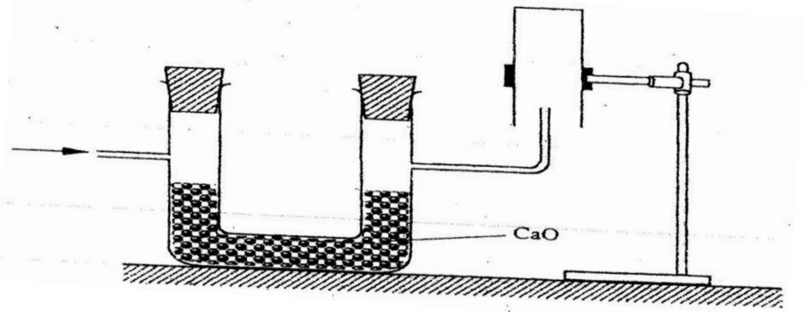
27. The diagram below shows a "Jiko" when in use. Study it and answer the questions that follow.



a) Identify the gas formed at region A. (1 mark)

b) State and explain the observation made at region B. (2 marks)

28. The set-up below was used to collect a dry sample of a gas.



Give two reasons why the set-up cannot be used to collect carbon (IV) oxide gas. (2 marks)

29. a). Describe how carbon (IV) oxide can be distinguished from carbon (II) oxide using calcium hydroxide (2 marks)

    b). State the role of carbon (IV) oxide in fire extinguishers(1 mark)

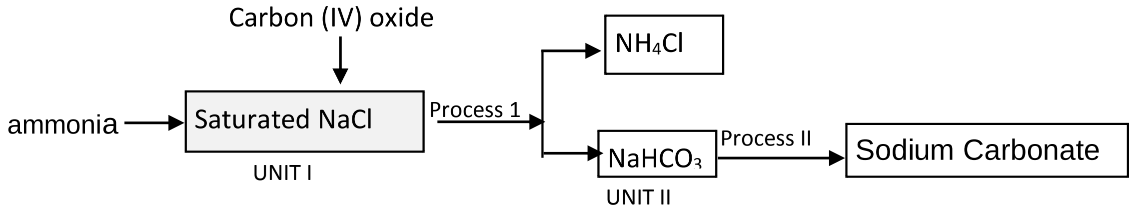
30. Carbon exists in different crystalline forms. Some of these forms were recently discovered in soot and are called fullerenes
    a) What name is given to different crystalline forms of the same element? (1 mark)
    b) Fullerenes dissolve in methylbenzene while the other forms of carbon do not. Given that soot is a mixture of fullerenes and other solid forms of carbon, describe how crystals of fullerenes can be obtained from soot. (3 marks)

31. When carbon (IV) oxide is bubbled in lime water, a white precipitate is observed, when excess carbon (IV) is use the white precipitate dissolves to form a colourless solution, when the colourless solution is boiled the white precipitate reappears. Explain these observations (3 marks)

32. When solid B_1 was heated, a gas which formed a white precipitate when passed through lime water was produced. The residue was dissolved in dilute nitric (V) acid to form a colourless solution B_2. When dilute hydrochloric acid was added to solution B_2 a white precipitate which dissolved on warming was formed. Write the formula of the;
    a) Cation in solid B_1 (1 mark)

    b) Anion in solid B_1 (1 mark)
33. When extinguishing fire caused by petrol, carbon (IV) oxide is used in preference to water explain (3 marks)

34. The Schematic diagram shows part of the Solvay process used for the manufacture of sodium carbonate



a) Explain how the sodium chloride required for this process is obtained from sea water (2 marks)

b) Two main reactions take place in UNIT 1.The first one is the formation of ammonium hydrogen carbonate

i. Write an equation for the reaction (1 mark)

ii. Write an equation for the second reaction (1mark)

c) State how the following are carried out (2 marks)

i. Process 1

ii. Process II

35. Carbon (II) oxide is described as a "silent killer"

a) State **one** physical property of carbon (II) oxide that makes it a "silent killer" (1 mark)

b) State and explain **one** chemical property that makes carbon (II) oxide poisonous to human beings (2 marks)

36. A water trough, aqueous sodium hydroxide, burning candle, watch class and a graduated gas jar were used in an experimental set up to determine the percentage of active part of air. Draw a labeled diagram of the set up at the end of the experiment. (3 marks)

37. Exhaust fumes of some cars contain carbon II oxide and other gases

a) Explain how carbon (II) oxide is formed in the internal combustion engines

i. (1 mark)

b) Name two gases other than carbon (II) oxide that are contained in exhaust fumes and are pollutants. (2 marks)

38. Graphite is one of the allotropes of carbon.

a) Name one other element which exhibits allotropy (1 mark)

b) Explain why graphite is used in the making of pencil leads. (2 marks)

39. Give two properties of carbon (IV) oxide which makes it suitable for use in fire extinguishers (2marks

40. a). Differentiate between allotropy and isotopy

b). Graphite is allotropy of carbon which conducts electricity though carbon is a non-metal. Explain

c). In terms of structure and bonding. Explain why diamond is used to make rock drills.

d). Carbon reacts with a hot concentrated Sulphuric (VI) acid

i. Give an equation for the reaction that occurs

ii. What property of carbon is displayed by this reaction

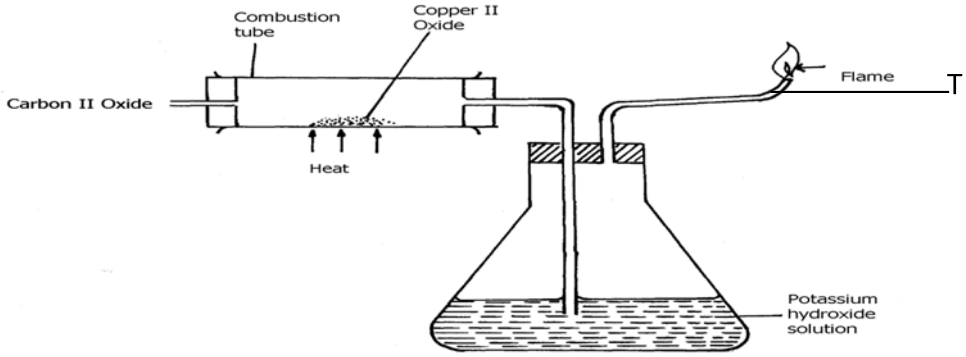
e). Sodium carbonate can be produced from trona as double salt

i. Write the formula of trona

ii. Why is trona a double salt

iii. How is trona converted to Na_2CO_3

41. Study the diagram below and answer the questions that follow



a) Explain the observation made in the combustion tube during the experiment.

b) Write an equation for the reaction that takes place in the combustion tube

c) What is responsible for the flame at the end of the tube marked T

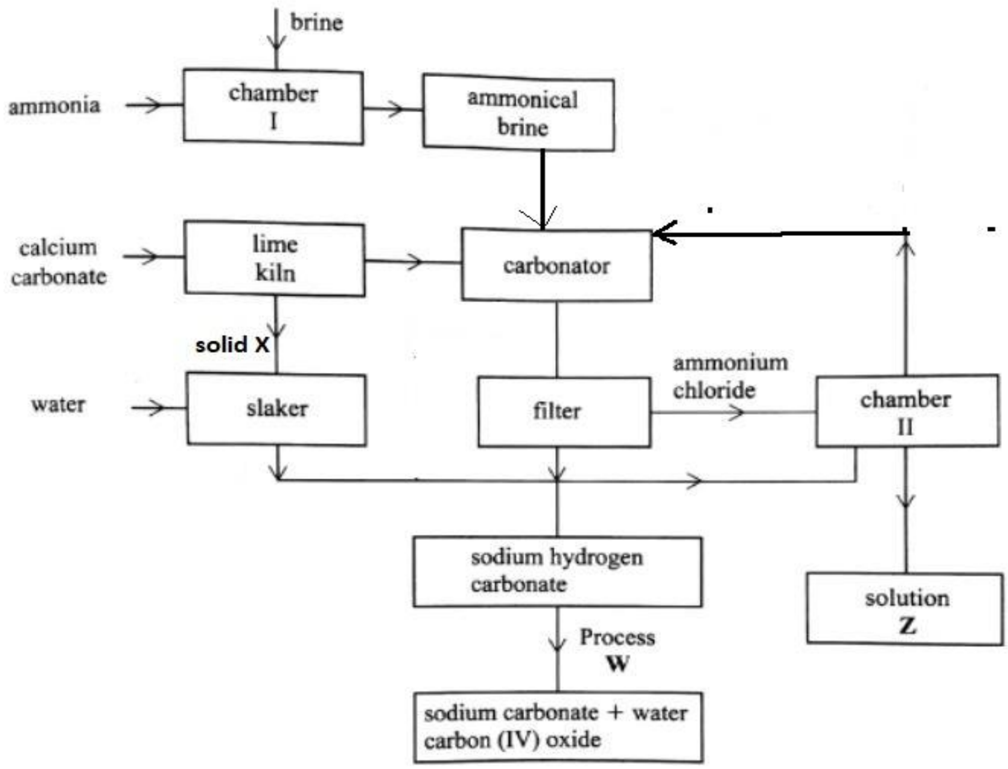
d) Explain the effect of increased carbon (IV) oxide in the atmosphere

42. a). in Kenya, sodium carbonate is extracted from trona at lake magadi.

i) Give the formula of trona.

ii) Name the process of extracting sodium carbonate from trona.

b) The flow chart in figure 5 summarizes the steps involved in the production of sodium carbonate. Use it to answer the questions that follow.



i) Name the process illustrated in figure 5.

ii) Identify the starting raw materials required in the production of sodium carbonate.

iii) Write equations for the two reactions that occur in the carbonator.

iv) Name two substances that are recycled.

v) Identify:

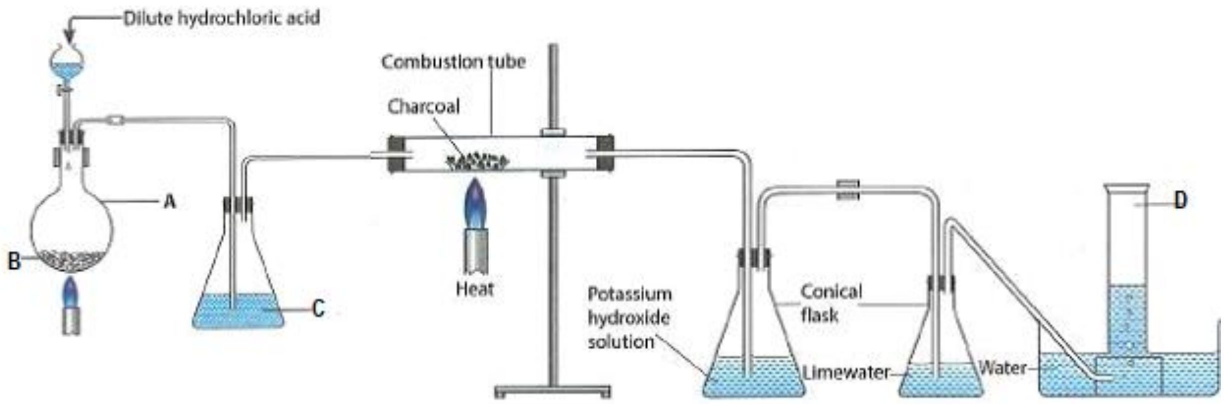
Solid X……………………………………………………

Process **W**……………………………………………………..

vi) Write an equation for the reaction that produces solution Z.

vii) Apart from softening hard water, state two other uses of sodium carbonate.

43. The diagram below was used to prepare dry carbon (IV) oxide and investigate some of its properties



I. Write chemical equations for the reactions that occur in flask A and in the combustion tube (2 marks)

II. Identify the following 2 marks

Reagent B

Substance C

And gas D

III. With the help of equation ,explain the use of potassium hydroxide in the set up 1mark

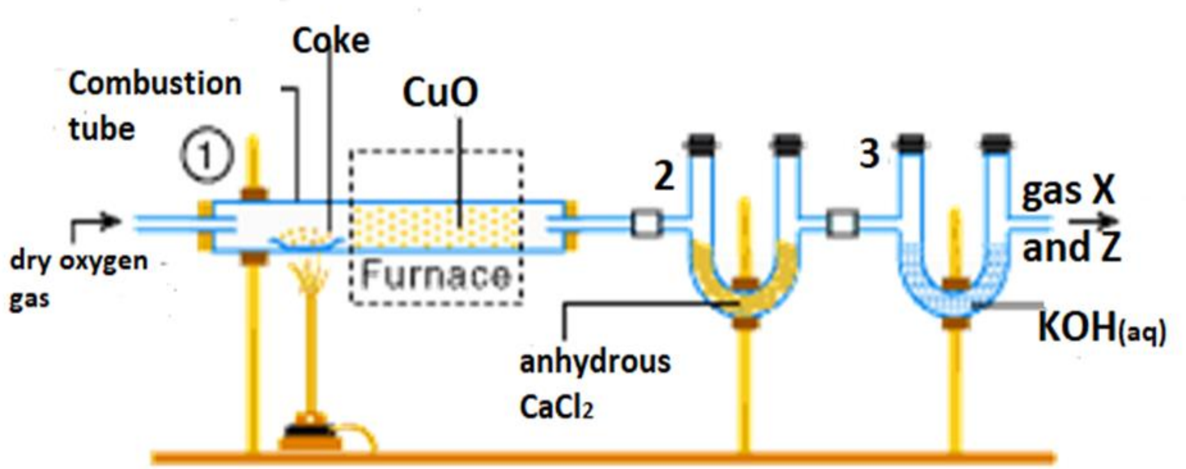
IV. Why is it not possible to collect pure and dry carbon (II) oxide using both downward delivery or upward delivery (1mark)

V. Give a the most suitable method that can be used to collect dry carbon (II) oxide (1 mark)

VI. Both Sulphur (IV) oxide and carbon (IV) oxide a denser than air, but only carbon (IV) oxide is used in fire extinguishers, Explain (2 marks)

VII. Carbon two oxide is referred as a silent killer, give two physical properties of carbon (II) oxie that make carbon (II) oxide a silent killer (2marks)

44. study The diagram below an answer the question that follow



a. Two gases are produced when coke reacts with oxygen name the two gases (2marks)

b. One of the gases is a reducing agent but very poisonous and can cause death, explain how it causes death(2marks)

c. Write equations for the reaction that occur in the combustion tube(2marks)

d. Write equations for the reaction that occur in the U-tube containing potassium hydroxide(2marks)

e. State the observations that occur in the furnace and U-tube 2 (2marks)

f. Suggest possible identities of gas W and X(2marks)

THIS IS THE LAST PRINTED PAGE