UPGRADE
CHEMISTRY
FORM 2

Comprehensive tutorial notes

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A.ATOMIC STRUCTURE

The atom is the smallest particle of an element that take part in a chemical reaction. The atom is made up of three subatomic particles:

(i)Protons
(ii)Electrons
(iii)Neutrons

(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 Electrons is negatively charged
2.Is found in fixed regions surrounding the centre of an atom called energy levels/orbitals.
3.It has a relative mass $^{1}/_{1840}$
4.The number of protons and electrons in a atom of an element is always equal

(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
4.The number of protons and neutrons in a atom of an element is its Mass number

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

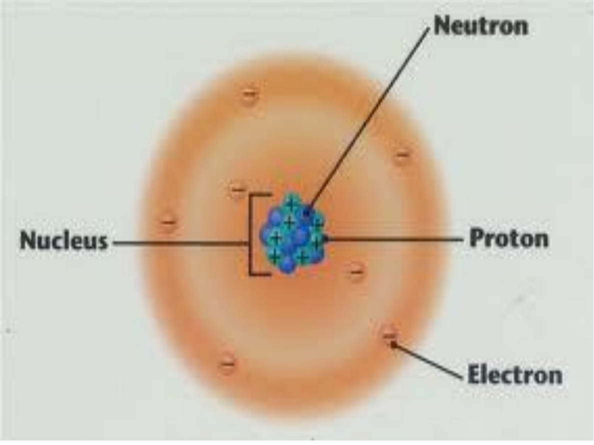
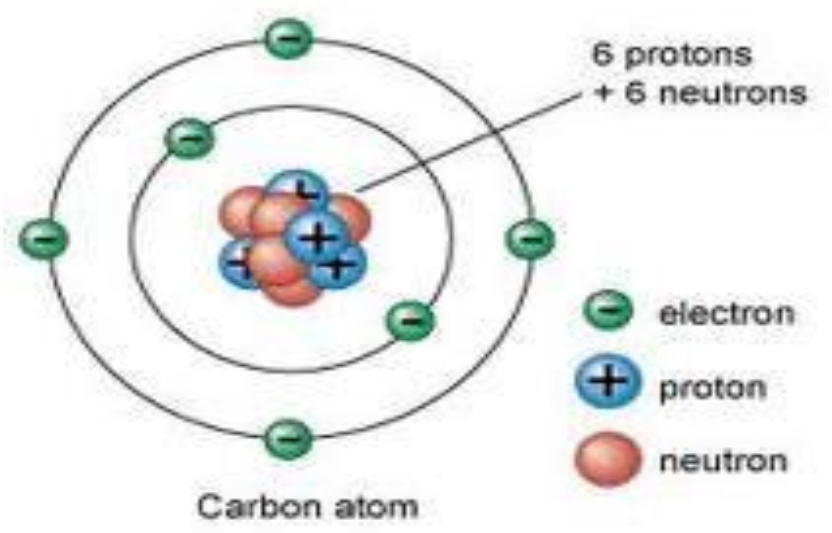


Diagram showing the relative positions of protons, electrons and neutrons in an atom of Carbon



The table below show atomic structure of the 1st twenty elements.

Element	**Symbol**	**Protons**	**Electrons**	**Neutrons**	**Atomic number**	**Mass number**
Hydrogen	**H**	**1**	**1**	**0**	**1**	**1**
Helium	**He**	**2**	**2**	**2**	**2**	**4**
Lithium	**Li**	**3**	**3**	**4**	**3**	**7**
Beryllium	**Be**	**4**	**4**	**5**	**4**	**9**
Boron	**B**	**5**	**5**	**6**	**5**	**11**
Carbon	**C**	**6**	**6**	**6**	**6**	**12**
Nitrogen	**N**	**7**	**7**	**7**	**7**	**14**
Oxygen	**O**	**8**	**8**	**8**	**8**	**16**
Fluorine	**F**	**9**	**9**	**10**	**9**	**19**
Neon	**Ne**	**10**	**10**	**10**	**10**	**20**
Sodium	**Na**	**11**	**11**	**12**	**11**	**23**
Magnesium	**Mg**	**12**	**12**	**12**	**12**	**24**
Aluminium	**Al**	**13**	**13**	**14**	**13**	**27**
Silicon	**Si**	**14**	**14**	**14**	**14**	**28**
Phosphorus	**P**	**15**	**15**	**16**	**15**	**31**
Sulphur	**S**	**16**	**16**	**16**	**16**	**32**
Chlorine	**Cl**	**17**	**17**	**18**	**17**	**35**
Argon	**Ar**	**18**	**18**	**22**	**18**	**40**
Potassium	**K**	**19**	**19**	**20**	**19**	**39**
Calcium	**Ca**	**20**	**20**	**20**	**20**	**40**

Most atoms of elements exist as **isotopes.**

Isotopes are atoms of the same **element**, having the same number of **protons**/atomic number but **different** number of **neutrons**/mass number.

By convention, isotopes are written with the mass number as superscript and the atomic number as subscript to the left of the chemical symbol of the element. i.e.

mass number → $^{m}_{n}X$ ← symbol of element
atomic number →

Below is the conventional method of writing the 1st twenty elements showing the mass numbers and atomic numbers;

$^{1}_{1}H$ $^{4}_{2}He$ $^{7}_{3}Li$ $^{9}_{4}Be$ $^{11}_{5}B$ $^{12}_{6}C$

$^{14}_{7}N$ $^{16}_{8}O$ $^{19}_{9}F$ $^{20}_{10}Ne$ $^{23}_{11}Na$ $^{24}_{12}Mg$

$^{27}_{13}Al$ $^{28}_{14}Si$ $^{31}_{15}P$ $^{32}_{16}S$ $^{35}_{17}Cl$ $^{40}_{18}Ar$

$^{39}_{19}K$ $^{40}_{20}C$

The table below shows some common **natural** isotopes of some elements

Element	Isotopes	Protons	Electrons	Neutrons	Atomic number	Mass number
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$(deuterium) $^{3}_{1}H$(Tritium)	1 1 1	1 1 1	0 2 3	1 1 1	1 2 3
Chlorine	$^{35}_{17}Cl$ $^{37}_{17}Cl$	17 17	17 17	18 20	17 17	35 37
Potassium	$^{39}_{19}K$ $^{40}_{19}K$ $^{41}_{19}K$	19 19 19	19 19 19	20 21 22	19 19 19	39 40 41
Oxygen	$^{16}_{8}O$ $^{18}_{8}O$	8 8	8 8	8 10	8 8	16 18
Uranium	$^{235}_{92}U$ $^{238}_{92}U$	92 92	92 92	143 146	92 92	235 238
Neon	$^{22}_{10}Ne$ $^{20}_{10}Ne$ $^{21}_{10}Ne$	10 10 10	10 10 10	12 10 11	10 10 10	22 20 21

The mass of an average atom is very small (10^{-22} g).Masses of atoms are therefore expressed in relation to a chosen element.

The atom recommended is ^{12}C isotope whose mass is arbitrarily assigned as 12.000 atomic mass units(a.m.u) .

All other atoms are compared to the mass of ^{12}C isotope to give the **relative at** The relative atomic mass(RAM) is therefore defined as "**the mass of average atom of an element compared to $^{1}/_{12}$ an atom of ^{12}C isotope whose mass is arbitrarily fixed as 12.000 atomic mass units(a.m.u)** " i.e;

$$\text{RAM} = \frac{\text{mass of atom of an element}}{{}^{1}/_{12} \text{ of one atom of } {}^{12}\text{C isotope}}$$

Accurate relative atomic masses (RAM) are got from the **mass spectrometer.** Mass spectrometer determines the **isotopes** of the element and their relative **abundance/availability.**

Using the relative abundances/availability of the isotopes, the relative atomic mass (RAM) can be determined /calculated as in the below examples.

a) Chlorine occurs as 75% $^{35}_{17}\text{Cl}$ and 25% $^{37}_{17}\text{Cl}$ isotopes. Calculate the relative atomic mass of Chlorine.

Working

100 atoms of chlorine contains 75 atoms of $^{35}_{17}\text{Cl}$ isotopes
100 atoms of chlorine contains 75 atoms of $^{37}_{17}\text{Cl}$ isotopes
Therefore;
RAM of chlorine = (75/100 x 35) + 25/100 x 37 = **35.5**
Note that:
Relative atomic mass has <u>no units</u>
More atoms of chlorine exist as $^{35}_{17}\text{Cl}$(75%) than as $^{37}_{17}\text{Cl}$(25%) therefore RAM is nearer to the more abundant isotope.

b) Calculate the relative atomic mass of potassium given that it exist as;
93.1% $^{39}_{19}\text{K}$, 0.01% $^{40}_{19}\text{K}$, 6.89% $^{41}_{19}\text{K}$,

Working

100 atoms of potassium contains 93.1 atoms of $^{39}_{19}\text{K}$ isotopes
100 atoms of potassium contains 0.01 atoms of $^{40}_{19}\text{K}$ isotopes
100 atoms of potassium contains 6.89 atoms of $^{41}_{19}\text{K}$ isotopes
Therefore;
RAM of potassium = (93.1/100 x39) + (0.01/100 x 40) +(6.89 /100 x 39)
=
Note that:
Relative atomic mass has <u>no units</u>
More atoms of potassium exist as $^{39}_{19}\text{K}$ (93.1%) therefore RAM is nearer to the more abundant $^{39}_{19}\text{K}$ isotope.

c) Calculate the relative atomic mass of Neon given that it exist as;
90.92% $^{20}_{10}\text{Ne}$, 0.26% $^{21}_{10}\text{Ne}$, 8.82% $^{22}_{10}\text{Ne}$,

Working

100 atoms of Neon contains 90.92 atoms of $^{20}_{10}\text{Ne}$ isotopes
100 atoms of Neon contains 0.26 atoms of $^{21}_{10}\text{Ne}$ isotopes
100 atoms of Neon contains 8.82 atoms of $^{22}_{10}\text{Ne}$ isotopes Therefore;
RAM of Neon = (90.92/100 x20) + (0.26/100 x 21) +(8.82 /100 x 22)

=

Note that:

Relative atomic mass has <u>no units</u>

More atoms of Neon exist as $^{20}_{10}Ne$ (90.92**%) therefore RAM is nearer to the more abundant $^{20}_{10}Ne$ isotope.**

d) Calculate the relative atomic mass of Argon given that it exist as;
90.92% $^{20}_{10}Ne$, 0.26% $^{21}_{10}Ne$, 8.82% $^{22}_{10}Ne$,
NB

The relative atomic mass is a measure of the masses of atoms. The <u>higher</u> the relative atomic mass, the <u>heavier</u> the atom.

Electrons are found in **energy levels/orbital.**
An energy level is a **fixed region** around/surrounding the nucleus of an atom occupied by electrons of the **same (potential) energy.**
By convention energy levels are named 1,2,3… **outwards** from the region **nearest** to nucleus.

Each energy level is occupied by a fixed number of electrons:

The **1st** energy level is occupied by a maximum of **two** electrons
The **2nd** energy level is occupied by a maximum of **eight** electrons
The **3rd** energy level is occupied by a maximum of **eight** electrons(or **eighteen** electrons if available)
The **4th** energy level is occupied by a maximum of **eight** electrons(or **eighteen or thirty two** electrons if available)

This arrangement of electrons in an atom is called **electron configuration** / **structure**.
By convention the electron configuration / structure of an atom of an element can be shown in form of a diagram using either cross(**x**) or dot(●) to

<u>Practice examples drawing electronic configurations</u>

a)$^{1}_{1}H$ has - in nucleus**1**proton and **0** neutrons
- 1 electron in the 1st energy levels thus:

Nucleus
Energy levels
Electrons(represented by cross(x)

Electronic structure of Hydrogen is thus: **1:**

b) $^{4}_{2}He$ has - in nucleus **2** proton and **2** neutrons - 2 electron in the 1st energy levels thus:

Nucleus
Energy levels

Electrons (represented by cross(x)

Electronic structure of Helium is thus: **2:**

c) $^{7}_{3}$Li has - in nucleus **3** proton and **4** neutrons
- **2** electron in the 1st energy levels
-**1** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Lithium is thus: **2:1**

d) $^{9}_{4}$Be has - in nucleus **4** proton and **5** neutrons
- **2** electron in the 1st energy levels
-**2** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Beryllium is thus: **2:2**

e) $^{11}_{5}$B has - in nucleus **5** proton and **6** neutrons
- **2** electron in the 1st energy levels
-**3** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Boron is thus: **2:3**

f) $^{12}_{6}$C has - in nucleus **6** proton and **6** neutrons
- **2** electron in the 1st energy levels
-**4** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Carbon is thus: **2:4**

g) $^{14}_{7}N$ has - in nucleus **7** proton and **7** neutrons
- **2** electron in the 1st energy levels
-**5** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Nitrogen is thus: **2:5**

h) $^{16}_{8}O$ has - in nucleus **8** proton and **8** neutrons
- **2** electron in the 1st energy levels
-**6** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Oxygen is thus: **2:6**

i) $^{19}_{9}F$ has - in nucleus **9** proton and **10** neutrons
- **2** electron in the 1st energy levels
-**7** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Fluorine is thus: **2:7**
i) $^{20}_{10}Ne$ has - in nucleus **10** proton and **10** neutrons
- **2** electron in the 1st energy levels

-8 electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

Electronic structure of Neon is thus: **2:8**

j) $^{23}_{11}Na$ has - in nucleus **11** proton and **12** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**1** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)

Electronic structure of Sodium is thus: **2:8:1**

k) $^{24}_{12}Mg$ has - in nucleus **12** proton and **12** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**2** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)

Electronic structure of Magnesium is thus: **2:8:2**
l) $^{27}_{13}Al$ has - in nucleus **13** proton and **14** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**3** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)

Electronic structure of Aluminium is thus: **2:8:3**

m) $^{28}_{14}Si$ has - in nucleus **14** proton and **14** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**4** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)
Electronic structure of Silicon is thus: **2:8:4**

n) $^{31}_{15}P$ has - in nucleus **14** proton and **15** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**5** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)
Electronic structure of Phosphorus is thus: **2:8:5**

o) $^{32}_{16}S$ has - in nucleus **16** proton and **16** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**6** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)
Electronic structure of Sulphur is thus: **2:8:6**

p) $^{35}_{17}Cl$ has - in nucleus **18** proton and **17** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**7** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)

Electronic structure of Chlorine is thus: **2:8:7**

p) $^{40}_{18}$Ar has - in nucleus **22** proton and **18** neutrons
- **2** electron in the 1st energy levels
- **8** electron in the 2nd energy levels
- **8** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)
Electronic structure of Argon is thus: **2:8:8**

q) $^{39}_{19}$K has - in nucleus **20** proton and **19** neutrons
- **2** electron in the 1st energy levels
- **8** electron in the 2nd energy levels
- **8** electron in the 3rd energy levels
- **1** electron in the 4th energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)
Electronic structure of Potassium is thus: **2:8:8:1**

r) $^{40}_{20}$Ca has - in nucleus **20** proton and **20** neutrons
- **2** electron in the 1st energy levels
- **8** electron in the 2nd energy levels
- **8** electron in the 3rd energy levels
- **2** electron in the 4th energy levels thus

Nucleus
Energy levels
Electrons (represented by dot(.)
Electronic structure of Calcium is thus: **2:8:8:2**

B.PERIODIC TABLE

There are over 100 elements so far discovered. Scientists have tried to group them together in a periodic table.

A periodic table is a horizontal and vertical arrangement of elements according to their atomic numbers.

This table was successfully arranged in 1913 by the British scientist Henry Moseley from the previous work of the Russian Scientist Dmitri Mendeleev.

The horizontal arrangement forms **period**. Atoms in the same period have the same the same number of energy levels in their electronic structure. i.e.

The number of **energy levels** in the electronic configuration of an element determine the **period** to which the element is in the periodic table.

e.g.

Which period of the periodic table are the following isotopes/elements/atoms?

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

   Electron structure 2:4 => 2 energy levels used thus **Period 2**

b) $^{23}_{11}Na$

   Electron structure 2:8:1 => 3 energy levels used thus **Period 3**

c) $^{39}_{19}K$

   Electron structure 2:8:8:1 => 4 energy levels used thus **Period 4**

d) $^{1}_{1}H$
   Electron structure 1: => 1 energy level used thus **Period 1**

The vertical arrangement of elements forms a **group.** Atoms in the same have the same the same group have the same number of outer energy level electrons as per their electronic structure. i.e.

The number of electrons in the outer energy level an element determine the **group** to which the element is ,in the periodic table.

a) $^{12}_{6}C$
   Electron structure 2:**4** => 4 electrons in outer energy level thus **Group IV**

b) $^{23}_{11}C$
   Electron structure 2:8:**1** => 1 electron in outer energy level thus **Group I**

c) $^{39}_{19}K$

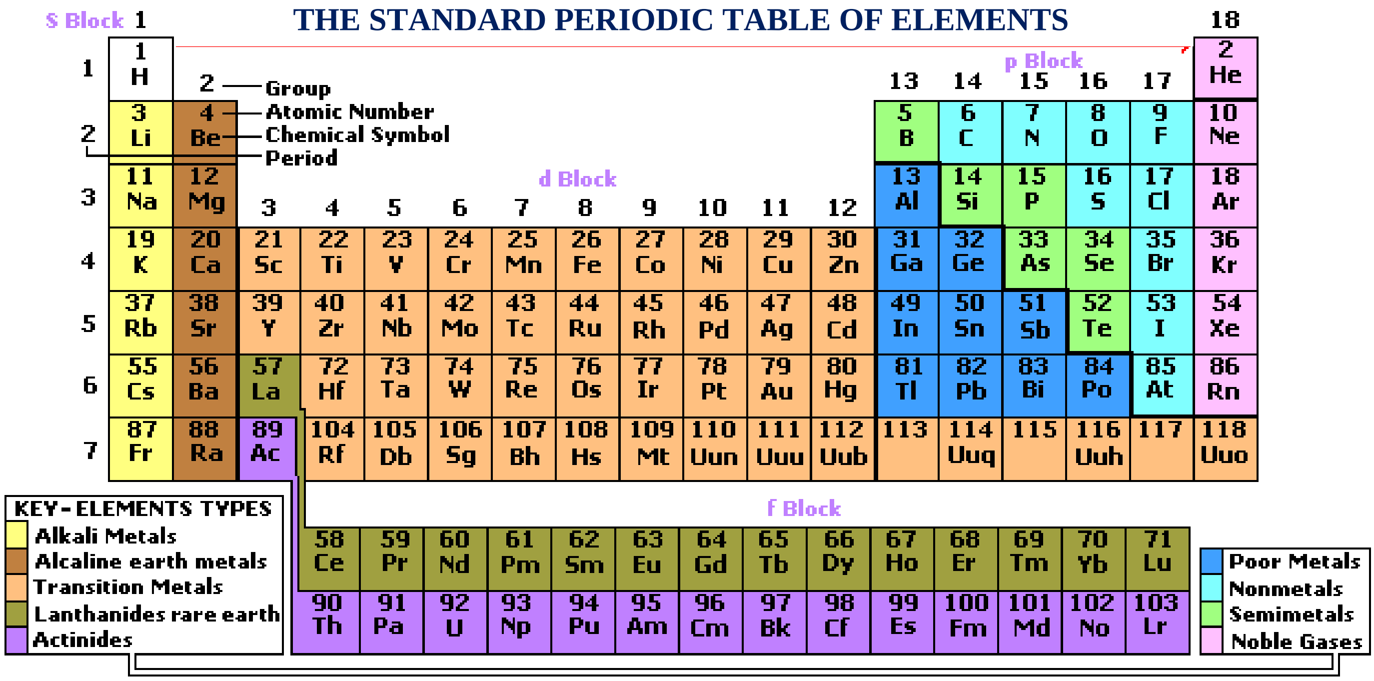
   Electron structure 2:8:8:**1**=>1 electron in outer energy level thus **Group I**

d) $^{1}_{1}H$
   Electron structure **1**: => 1 electron in outer energy level thus **Group I**

By convention;

(i)**Periods** are named using English numerals **1,2,3,4**,…
(ii)**Groups** are named using Roman numerals **I,II,III,IV**,…

There are eighteen groups in a standard periodic table.
There are seven periods in a standard periodic table.



THE STANDARD PERIODIC TABLE OF ELEMENTS

s Block

p Block

d Block

2 — Group
4 — Atomic Number
Be — Chemical Symbol
2 — Period

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub	113	114 Uuq	115	116 Uuh	117	118 Uuo

f Block

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

KEY - ELEMENTS TYPES
Alkali Metals
Alcaline earth metals
Transition Metals
Lanthanides rare earth
Actinides

Poor Metals
Nonmetals
Semimetals
Noble Gases

When an atom has maximum number of electrons in its outer energy level, it is said to be **stable.**

When an atom has no maximum number of electrons in its outer energy level, it is said to be **unstable.**

All stable atoms are in group **8/18** of the periodic table. All other elements are unstable.

All unstable atoms/isotopes try to be stable through chemical reactions. A chemical reaction involves gaining or losing outer electrons (electron transfer) .When electron transfer take place, an ion is formed.

An ion is formed when an unstable atom gain or donate electrons in its outer energy level inorder to be stable. Whether an atom gain or donate electrons depend on the relative energy required to donate or gain extra electrons i.e.

Examples

1. $^{19}_{9}$F has electronic structure/configuration 2:7.
   It can donate the seven outer electrons to have stable electronic structure/configuration 2:.
   It can gain one extra electron to have stable electronic structure/configuration 2:8. Gaining requires less energy, and thus Fluorine reacts by gaining one extra electrons.
2. $^{23}_{13}$Al has electronic structure/configuration 2:8:3
   It can donate the three outer electrons to have stable electronic structure/configuration 2:8.

It can gain five extra electrons to have stable electronic structure/configuration 2:8:8. Donating requires less energy, and thus Aluminium reacts by donating its three outer electrons.

Elements with **less** than four electrons in the outer energy level donates /lose the outer electrons to be stable and form a positively charged ion called **cation.**
A cation therefore has more protons(positive charge) than electrons(negative charge)
Generally metals usually form cation
Elements with **more** than four electrons in the outer energy level gain /acquire extra electrons in the outer energy level to be stable and form a negatively charged ion called **anion.**
An anion therefore has less protons(positive charge) than electrons(negative charge)
Generally non metals usually form anion. Except Hydrogen
The charge carried by an ion is equal to the number of electrons gained/acquired or donated/lost.

Examples of ion formation

1. $^{1}_{1}H$
H -> H^+ + e
(atom) (monovalent cation) (electrons donated/lost)
Electronic configuration 1: (No electrons remains)

2. $^{27}_{13}Al$
Al -> Al^{3+} + 3e
(atom) (trivalent cation) (3 electrons donated/lost)
Electron structure 2:8:3 (unstable) 2:8 (stable)

3. $^{23}_{11}Na$
Na -> Na^+ + e
(atom) (cation) (1 electrons donated/lost)
Electron structure 2:8:1 (unstable) 2:8 (stable)

4. $^{24}_{12}Mg$
Mg -> Mg^{2+} + 2e
(atom) (cation) (2 electrons donated/lost)
Electron structure 2:8:1 (unstable) 2:8 (stable)

5. $^{16}_{8}O$
O + 2e -> O^{2-}
(atom) (2 electrons gained/acquired) (anion)

Electron structure 2:6 (unstable) 2:8 (stable)

6. $^{14}_{7}N$

N + 3e -> N^{3-}
(atom) (3 electrons gained/acquired) (anion)
Electron structure 2:5 (unstable) 2:8 (stable)

7. $^{31}_{15}P$

P + 3e -> P^{3-}
(atom) (3 electrons gained/acquired) (anion)
Electron structure 2:5 (unstable) 2:8 (stable)

8. $^{19}_{9}F$

F + e -> F^{-}
(atom) (1 electrons gained/acquired) (anion)
Electron structure 2:7 (unstable) 2:8 (stable)

9. $^{35}_{17}Cl$

Cl + e -> Cl^{-}
(atom) (1 electrons gained/acquired) (anion)
Electron structure 2:8:7 (unstable) 2:8:8 (stable)

3. $^{39}_{19}K$

K -> K^{+} + e
(atom) (cation) (1 electrons donated/lost)
Electron structure 2:8:8:1 (unstable) 2:8:8 (stable)

When an element donate/loses its outer electrons ,the process is called oxidation. When an element acquires/gains extra electrons in its outer energy level,the process is called reduction.The charge carried by an atom, cation or anion is its oxidation state.

Table showing the oxidation states of some isotopes

Element	Symbol of element / isotopes	Charge of ion	Oxidation state
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$(deuterium) $^{3}_{1}H$(Tritium)	H^{+} H^{+} H^{+}	+1 +1 +1

Chlorine	$^{35}_{17}Cl$ $^{37}_{17}Cl$	Cl^- Cl^-	-1 -1
Potassium	$^{39}_{19}K$ $^{40}_{19}K$ $^{41}_{19}K$	K^+ K^+ K^+	+1 +1 +1
Oxygen	$^{16}_{8}O$ $^{18}_{8}O$	O^{2-} O^{2-}	-2 -2
Magnesium	$^{24}_{12}Mg$	Mg^{2+}	+2
sodium	$^{23}_{11}Na$	Na^+	+1
Copper	Cu	Cu^+ Cu^{2+}	+1 +2
Iron		Fe^{2+} Fe^{3+}	+2 +3
Lead		Pb^{2+} Pb^{4+}	+2 +4
Manganese		Mn^{2+} Mn^{7+}	+2 +7
Chromium		Cr^{3+} Cr^{6+}	+3 +6
Sulphur		S^{4+} S^{6+}	+4 +6
Carbon		C^{2+} C^{4+}	+2 +4

Note :

Some elements can exist in **more** than one oxidation state.They are said to have variable oxidation state.

Roman capital numeral is used to indicate the oxidation state of an element with a variable oxidation state in a compound.

Examples:

(i) Copper (I) means Cu^+ as in Copper(I)oxide
(ii) Copper (II) means Cu^{2+} as in Copper(II)oxide
(iii) Iron (II) means Fe^{2+} as in Iron(II)sulphide
(iv) Iron (III) means Fe^{3+} as in Iron(III)chloride
(iv) Sulphur(VI)mean S^{6+} as in Iron(III)sulphate(VI)
(v) Sulphur(VI)mean S^{6+} as in sulphur(VI)oxide
(vi) Sulphur(IV)mean S^{4+} as in sulphur(IV)oxide

(vii) Sulphur(IV)mean S^{4+} as in sodium sulphate(IV)
(ix) Carbon(IV)mean C^{4+} as in carbon(IV)oxide
(x) Carbon(IV)mean C^{4+} as in Lead(II)carbonate(IV)
(xi) Carbon(II)mean C^{2+} as in carbon(II)oxide
(xii) Manganese(IV)mean Mn^{4+} as in Manganese(IV)oxide

A compound is a combination of two or more elements in fixed proportions. The ratio of the atoms making a compound is called the chemical formulae. Elements combine together to form a compound depending on their combining power.

The combining power of atoms in an element is called Valency.Valency of an element is equal to the **number** of:

(i)hydrogen atoms that an atom of element can combine with or displace.
(ii)electrons gained /acquired in outer energy level by non metals to be stable/attain duplet/octet.
(iii)electrons donated/lost by outer energy level of metals to be stable/attain octet/duplet.
(iv)charges carried by ions/cations/ions

Group of atoms that react as a unit during chemical reactions are called **radicals**.Elements with variable oxidation state also have more than one valency.

Table showing the valency of common radicals.

Radical name	**Chemical formulae**	**Combining power / Valency**
Ammonium	NH_4^{+}	1
Hydroxide	OH^{-}	1
Nitrate(V)	NO_3^{-}	1
Hydrogen carbonate	HCO_3^{-}	1
Hydrogen sulphate(VI)	HSO_4^{-}	1
Hydrogen sulphate(IV)	HSO_3^{-}	1
Manganate(VII)	MnO_4^{-}	1
Chromate(VI)	CrO_4^{2-}	2
Dichromate(VI)	$Cr_2O_7^{2-}$	2
Sulphate(VI)	SO_4^{2-}	2
Sulphate(IV)	SO_3^{2-}	2
Carbonate(IV)	CO_3^{2-}	2
Phosphate(V)	PO_4^{2-}	3

Table showing the valency of some common metal and non metals

Element/metal	Valency	**Element/non metal**	**Valency**
Hydrogen	1	**Florine**	**1**
Lithium	1	**Chlorine**	**1**
Beryllium	2	**Bromine**	**1**

Boron	3	**Iodine**	**1**
Sodium	1	**Carbon**	**4**
Magnesium	2	**Nitrogen**	**3**
Aluminium	3	**Oxygen**	**2**
Potassium	1	**Phosphorus**	**3**
Calcium	2		
Zinc	2		
Barium	2		
Mercury	2		
Iron	2 and 3		
Copper	1 and 2		
Manganese	2 and 4		
Lead	2 and 4		

From the valency of elements , the chemical formular of a compound can be derived using the following procedure:

(i)Identify the elements and radicals making the compound

(ii)Write the symbol/formular of the elements making the compound starting with the metallic element

(iii)Assign the valency of each element /radical as superscript.

(iv)Interchange/exchange the valencies of each element as subscript.

(v)Divide by the smallest/lowest valency to derive the smallest whole number ratios Ignore a valency of 1.

This is the chemical formula.

<u>Practice examples</u>

Write the chemical formula of

(a)Aluminium oxide

Elements making compound	Aluminium	Oxygen
Symbol of elements/radicals in compound	Al	O
Assign valencies as superscript	Al^3	O^2
Exchange/Interchange the valencies as subscript	Al_2	O_3
Divide by smallest valency to get whole number	-	-

Chemical formula of Aluminium oxide is thus**:** **<u>Al_2O_3</u>**

This means:2atoms of Aluminium combine with 3 atoms of Oxygen

(b)Sodium oxide

Elements making compound	Sodium	Oxygen
Symbol of elements/radicals in compound	Na	O
Assign valencies as superscript	Na^1	O^2
Exchange/Interchange the valencies as subscript	Na_2	O_1

Divide by smallest valency to get whole number	-	-

Chemical formula of Sodium oxide is thus**: $\underline{Na_2 O}$**
This means:2atoms of Sodium combine with 1 atom of Oxygen

(c)Calcium oxide

Elements making compound	Calcium	Oxygen
Symbol of elements/radicals in compound	Ca	O
Assign valencies as superscript	Ca^2	O^2
Exchange/Interchange the valencies as subscript	Ca_2	O_2
Divide by two to get smallest whole number ratio	Ca_1	O_1

Chemical formula of Calcium oxide is thus**: $\underline{CaO}$**
This means:1 atom of calcium combine with 1 atom of Oxygen.

(d)Lead(IV)oxide

Elements making compound	Lead	Oxygen
Symbol of elements/radicals in compound	Pb	O
Assign valencies as superscript	Pb^4	O^2
Exchange/Interchange the valencies as subscript	Pb_2	O_4
Divide by two to get smallest whole number ratio	Pb_1	O_2

Chemical formula of Lead(IV) oxide is thus**: $\underline{PbO_2}$**
This means:1 atom of lead combine with 2 atoms of Oxygen.

(e)Lead(II)oxide

Elements making compound	Lead	Oxygen
Symbol of elements/radicals in compound	Pb	O
Assign valencies as superscript	Pb^2	O^2
Exchange/Interchange the valencies as subscript	Pb_2	O_2
Divide by two to get smallest whole number ratio	Pb_1	O_1

Chemical formula of Lead(II) oxide is thus**: $\underline{PbO}$**
This means:1 atom of lead combine with 1 atom of Oxygen.

(e)Iron(III)oxide

Elements making compound	Iron	Oxygen
Symbol of elements/radicals in compound	Fe	O
Assign valencies as superscript	Fe^3	O^2
Exchange/Interchange the valencies as subscript	Fe_2	O_3

Divide by two to get smallest whole number ratio	-	-

Chemical formula of Iron(III) oxide is thus**: Fe_2O_3**
This means:2 atom of lead combine with 3 atom of Oxygen.

(f)Iron(II)sulphate(VI)

Elements making compound	Iron	sulphate(VI)
Symbol of elements/radicals in compound	Fe	SO_4
Assign valencies as superscript	Fe^2	$SO_4{}^2$
Exchange/Interchange the valencies as subscript	Fe_2	$SO_4{}_2$
Divide by two to get smallest whole number ratio	Fe_1	$SO_4{}_1$

Chemical formula of Iron(II) sulphate(VI) is thus**: $FeSO_4$**
This means:1 atom of Iron combine with 1 sulphate(VI) radical.

(g)Copper(II)sulphate(VI)

Elements making compound	Copper	sulphate(VI)
Symbol of elements/radicals in compound	Cu	SO_4
Assign valencies as superscript	Cu^2	$SO_4{}^2$
Exchange/Interchange the valencies as subscript	Cu_2	$SO_4{}_2$
Divide by two to get smallest whole number ratio	Cu_1	$SO_4{}_1$

Chemical formula of Cu(II)sulphate(VI) is thus**: $CuSO_4$**
This means:1 atom of Copper combine with 1 sulphate(VI) radical.

(h)Aluminium sulphate(VI)

Elements making compound	Aluminium	sulphate(VI)
Symbol of elements/radicals in compound	Al	SO_4
Assign valencies as superscript	Al^3	$SO_4{}^2$
Exchange/Interchange the valencies as subscript	Al_2	$SO_4{}_3$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Aluminium sulphate(VI) is thus**: $Al_2(SO_4)_3$**
This means:2 atom of Aluminium combine with 3 sulphate(VI) radical.

(i)Aluminium nitrate(V)

Elements making compound	Aluminium	nitrate(V)
Symbol of elements/radicals in compound	Al	NO_3
Assign valencies as superscript	Al^3	$NO_3{}^1$
Exchange/Interchange the valencies as subscript	Al_1	$NO_3{}_3$

Divide by two to get smallest whole number ratio	-	-

Chemical formula of Aluminium sulphate(VI) is thus: **$Al (NO_3)_3$**
This means:1 atom of Aluminium combine with 3 nitrate(V) radical.

(j)Potassium manganate(VII)

Elements making compound	Potassium	manganate(VII)
Symbol of elements/radicals in compound	K	MnO_4
Assign valencies as superscript	K^1	$MnO_4{}^1$
Exchange/Interchange the valencies as subscript	K_1	$MnO_4{}_1$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Potassium manganate(VII) is thus: **$KMnO_4$**
This means:1 atom of Potassium combine with 4 manganate(VII) radical.

(k)Sodium dichromate(VI)

Elements making compound	Sodium	dichromate(VI)
Symbol of elements/radicals in compound	Na	Cr_2O_7
Assign valencies as superscript	Na^1	$Cr_2O_7{}^2$
Exchange/Interchange the valencies as subscript	Na_2	$Cr_2O_7{}_1$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Sodium dichromate(VI) is thus: **$Na_2\ Cr_2O_7$**
This means:2 atom of Sodium combine with 1 dichromate(VI) radical.

(l)Calcium hydrogen carbonate

Elements making compound	Calcium	Hydrogen carbonate
Symbol of elements/radicals in compound	Ca	CO_3
Assign valencies as superscript	Ca^2	$HCO_3{}^1$
Exchange/Interchange the valencies as subscript	Ca_1	$HCO_3{}_2$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Calcium hydrogen carbonate is thus: **$Ca(HCO_3)_2$**
This means:1 atom of Calcium combine with 2 hydrogen carbonate radical.

(l)Magnesium hydrogen sulphate(VI)

Elements making compound	Magnesium	Hydrogen sulphate(VI)
Symbol of elements/radicals in compound	Mg	HSO_4

Assign valencies as superscript	Mg^{2}	$HSO_4{}^{1}$
Exchange/Interchange the valencies as subscript	Mg_1	$HSO_4{}_{2}$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Magnesium hydrogen sulphate(VI) is thus**:** **$Mg(HSO_4)_2$**
This means:1 atom of Magnesium combine with 2 hydrogen sulphate(VI) radical.

Compounds are formed from chemical reactions. A chemical reaction is formed when atoms of the reactants break free to bond again and form products. A chemical reaction is a statement showing the movement of reactants to form products. The following procedure is used in writing a chemical equations:

1. Write the word equation
2. Write the correct chemical formula for each of the reactants and products
3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.
4. Multiply the chemical formula containing the unbalanced atoms with the lowest common multiple if the number of atoms on one side is not equal. This is called **balancing.**
Do not change the chemical formula of the products/reactants.
5. Assign in brackets, the physical state/state symbols of the reactants and products after each chemical formula as:
    (i) **(s)** for solids
    (ii) **(l)** for liquids
    (iii) **(g)** for gas
    (iv) **(aq)** for aqueous/dissolved in water to make a solution.

Practice examples

Write a balanced chemical equation for the following

(a) Hydrogen gas is prepared from reacting Zinc granules with dilute hydrochloric acid.
Procedure

1. Write the word equation
Zinc + Hydrochloric acid -> Zinc chloride + hydrogen gas

2. Write the correct chemical formula for each of the reactants and products
$Zn + HCl \rightarrow ZnCl_2 + H_2$

3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

Number of atoms of Zn on the reactant side is equal to product side

One atom of H in HCl on the reactant side is not equal to two atoms in H_2 on product side.

One atom of Cl in HCl on the reactant side is not equal to two atoms in $ZnCl_2$ on product side.

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

Multiply HCl by "2" to get "2" Hydrogen and "2" Chlorine on product and reactant side.

$Zn + \underline{2}HCl \rightarrow ZnCl_2 + H_2$

5. Assign in brackets, the physical state/state symbols .

$Zn(s) + \underline{2}HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

(b) Oxygen gas is prepared from decomposition of Hydrogen peroxide solution to water

Procedure

1. Write the word equation

Hydrogen peroxide -> Water + oxygen gas

2. Write the correct chemical formula for each of the reactants and products

$H_2O_2 \rightarrow H_2O + O_2$

3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

Number of atoms of H on the reactant side is <u>equal</u> to product side

Two atom of O in H_2O_2 on the reactant side is <u>not equal</u> to three atoms (one in H_2O and two in O_2) on product side.

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

Multiply H_2O_2 by "2" to get "4" Hydrogen and "4" Oxygen on reactants

Multiply H_2O by "2" to get "4" Hydrogen and "2" Oxygen on product side

When the "2" Oxygen in O_2 and the"2" in H_2O are added on product side they are equal to the"4" Oxygen on reactants side.

$\underline{2}H_2O_2 \rightarrow \underline{2}H_2O + O_2$

5. Assign in brackets, the physical state/state symbols .

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

(c) Chlorine gas is prepared from Potassium manganate(VII) reacting with hydrochloric acid to form potassium chloride solution, manganese(II) chloride solution,water and chlorine gas.

Procedure

1. Write the word equation

Potassium manganate(VII) + Hydrochloric acid ->
potassium chloride + manganese(II) chloride + chlorine **+**water

2. Write the correct chemical formula for each of the reactants and products

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

3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

Number of atoms of K and Mn on the reactant side is <u>equal</u> to product side

Two atom of H in H_2O on the product side is <u>not equal</u> to one atom on reactant side.

Four atom of O in $KMnO_4$ is not equal to one in H_2O

One atom of Cl in HCl on reactant side is not equal to three (one in H_2O and two in Cl_2)

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

Multiply HCl by “16” to get “16” Hydrogen and “16” Chlorine on reactants

Multiply $KMnO_4$ by “2” to get “2” Potassium and “2” manganese, “2 x4 =8” Oxygen on reactant side.

Balance the product side to get:

$$\underline{2}\,KMnO_4 + \underline{16}\,HCl \rightarrow \underline{2}\,KCl + \underline{2}\,MnCl_2 + \underline{8}\,H_2O + \underline{5}\,Cl_2$$

5. Assign in brackets, the physical state/state symbols .

$$\underline{2}KMnO_4(s) + 16\,HCl(aq) \rightarrow 2\,KCl\,(aq) + 2MnCl_2(aq) + 8\,H_2O(l) + 5\,Cl_2(g)$$

(d)Carbon(IV)oxide gas is prepared from Calcium carbonate reacting with hydrochloric acid to form calcium chloride solution, water and carbon(IV)oxide gas.

Procedure

1. Write the word equation

Calcium carbonate + Hydrochloric acid ->
calcium chloride solution+ water +carbon(IV)oxide

2. Write the correct chemical formula for each of the reactants and products

$$CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$$

3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

5. Assign in brackets, the physical state/state symbols .

$$CaCO_3(s) + 2\,HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

(d)Sodium hydroxide solution neutralizes hydrochloric acid to form salt and water.

$$NaOH(aq) + HCl(aq) \rightarrow NaCl\,(aq) + H_2O(l)$$

(e)Sodium reacts with water to form sodium hydroxide and hydrogen gas.

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

(f)Calcium reacts withwater to form calcium hydroxide and hydrogen gas

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

(g)Copper(II)Oxide solid reacts with dilute hydrochloric acid to form copper(II)chloride and water.

$$\mathbf{CuO(s) + \underline{2}HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)}$$

(h)Hydrogen sulphide reacts with Oxygen to form sulphur(IV)Oxide and water.

$$\mathbf{\underline{2}H_2S(g) + \underline{3}O_2(g) \rightarrow \underline{2}SO_2(g) + \underline{2}H_2O(l)}$$

(i)Magnesium reacts with steam to form Magnesium Oxide and Hydrogen gas.

$$\mathbf{Mg(s) + \underline{2}H_2O(g) \rightarrow MgO(s) + H_2(g)}$$

(j)Ethane(C_2H_6) gas burns in air to form Carbon(IV)Oxide and water.

$$\mathbf{\underline{2}C_2H_6(g) + \underline{7}O_2(g) \rightarrow \underline{4}CO_2(g) + \underline{6}H_2O(l)}$$

(k)Ethene(C_2H_4) gas burns in air to form Carbon(IV)Oxide and water.

$$\mathbf{C_2H_4(g) + \underline{3}O_2(g) \rightarrow \underline{2}CO_2(g) + \underline{2}H_2O(l)}$$

(l)Ethyne(C_2H_2) gas burns in air to form Carbon(IV)Oxide and water.

$$2\mathbf{C_2H_2(g) + \underline{5}O_2(g) \rightarrow \underline{4}CO_2(g) + \underline{2}H_2O(l)}$$

C.PERIODICITY OF CHEMICAL FAMILES/DOWN THE GROUP.

The number of valence electrons and the number of occupied energy levels in an atom of an element determine the position of an element in the periodic table.i.e

The number of occupied energy levels determine the Period and the valence electrons determine the Group.

Elements in the same group have similar physical and chemical properties. The trends in physical and chemical properties of elements in the same group vary down the group. Elements in the same group thus constitute a chemical family.

(a) Group I elements: Alkali metals

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

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

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

Atomic and ionic radius

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

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

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

Table showing the atomic and ionic radius of some alkali metals

Element	Symbol	Atomic number	Atomic radius(nM)	Ionic radius(nM)
Lithium	Li	3	0.133	0.060

Sodium	Na	11	0.157	0.095
Potassium	K	19	0.203	0.133

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

Electropositivity

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

Ionization energy

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

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

Physical properties

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

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

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

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

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

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

Chemical properties

(i)Reaction with air/oxygen

On exposure to air, alkali metals reacts with the elements in the air.

Example

On exposure to air, Sodium first reacts with Oxygen to form sodium oxide.

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

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

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

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

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

(ii)Burning in air/oxygen

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

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

Sodium burns in air with a **yellow** flame to form sodium oxide

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

Sodium burns in oxygen with a **yellow** flame to form sodium peroxide

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

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

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

(iii) Reaction with water:

Experiment

Measure 500 cm3 of water into a beaker.

Put three drops of phenolphthalein indicator.

Put about 0.5g of Lithium metal into the beaker.

Determine the pH of final product

Repeat the experiment using about 0.1 g of Sodium and Potassium.

Caution: Keep a distance

Observations

Alkali metal	Observations	Comparative speed/rate

		of the reaction
Lithium	-Metal floats in water -rapid effervescence/fizzing/bubbling -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Moderately vigorous
Sodium	-Metal floats in water -very rapid effervescence /fizzing /bubbling -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Very vigorous
Potassium	-Metal floats in water -explosive effervescence /fizzing /bubbling -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Explosive/burst into flames

Explanation

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

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

Chemical equations

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

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

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

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

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

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

Reactivity increase down the group

(iv) Reaction with chlorine:

Experiment

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

Repeat with about 0.5g of Lithium.

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

Observation

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

Lithium metal continues to burn with a crimson flame forming white solid / fumes.

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

Chemical equations

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

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

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

$2Rb(s) + Cl_2(g) \rightarrow 2RbCl(s)$

$2Cs(s) + Cl_2(g) \rightarrow 2CsCl(s)$

$2Fr(s) + Cl_2(g) \rightarrow 2FrCl(s)$

Reactivity increase down the group

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

	Lithium	sodium	Potassium
Hydroxide	LiOH	NaOH	KOH
Oxide	Li_2O	Na_2O	K_2O
Sulphide	Li_2S	Na_2S	K_2S
Chloride	LiCl	NaCl	KCl
Carbonate	Li_2CO_3	Na_2CO_3	K_2CO_3
Nitrate(V)	$LiNO_3$	$NaNO_3$	KNO_3
Nitrate(III)	-	$NaNO_2$	KNO_2
Sulphate(VI)	Li_2SO_4	Na_2SO_4	K_2SO_4
Sulphate(IV)	-	Na_2SO_3	K_2SO_3
Hydrogen carbonate	-	$NaHCO_3$	$KHCO_3$
Hydrogen sulphate(VI)	-	$NaHSO_4$	$KHSO_4$

Hydrogen sulphate(IV)	-	$NaHSO_3$	$KHSO_3$
Phosphate	-	Na_3PO_4	K_3PO_4
Manganate(VI)	-	$NaMnO_4$	$KMnO_4$
Dichromate(VI)	-	$Na_2Cr_2O_7$	$K_2Cr_2O_7$
Chromate(VI)	-	Na_2CrO_4	K_2CrO_4

Some **uses** of alkali metals include:

(i)Sodium is used in making sodium cyanide for extracting gold from gold ore.
(ii)Sodium chloride is used in seasoning food.
(iii)Molten mixture of sodium and potassium is used as coolant in nuclear reactors.
(iv)Sodium is used in making sodium hydroxide used in making soapy and soapless detergents.
(v)Sodium is used as a reducing agent for the extraction of titanium from Titanium(IV)chloride.
(vi)Lithium is used in making special high strength glasses
(vii)Lithium compounds are used to make dry cells in mobile phones and computer laptops.

Group II elements: Alkaline earth metals

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

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

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

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

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

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

Element	Symbol	Atomic	Atomic radius(nM)	Ionic radius(nM)

		number		
Beryllium	Be	4	0.089	0.031
Magnesium	Mg	12	0.136	0.065
Calcium	Ca	20	0.174	0.099

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

Electropositivity

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

Ionization energy

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

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

Physical properties

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

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

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

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

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

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

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

Calcium is a better conductor than magnesium.

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

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<u>Summary of some physical properties of the 1st three alkaline earth metals</u>

Alkaline earth metal	Appearance	Ease of cutting	Melting point (^{o}C)	Boiling point (^{o}C)	Conduct-ivity	1st ionization energy	2nd ionization energy

Beryllium	Shiny grey	Not easy	1280	3450	Good	900	1800
Magnesium	Shiny grey	Not Easy	650	1110	Good	736	1450
calcium	Shiny grey	Not easy	850	1140	Good	590	970

Chemical properties

(i)Reaction with air/oxygen

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

Example

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

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

(ii)Burning in air/oxygen

Experiment

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

Caution: Do not look directly at the flame

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

Repeat with Calcium

Observations

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

Explanation

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

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

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

Magnesium oxide dissolves in water to form magnesium hydroxide.

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

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

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

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

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

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

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

Calcium oxide dissolves in water to form calcium hydroxide.

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

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

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

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

(iii)Reaction with water

Experiment

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

Observations

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

Explanations

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

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

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

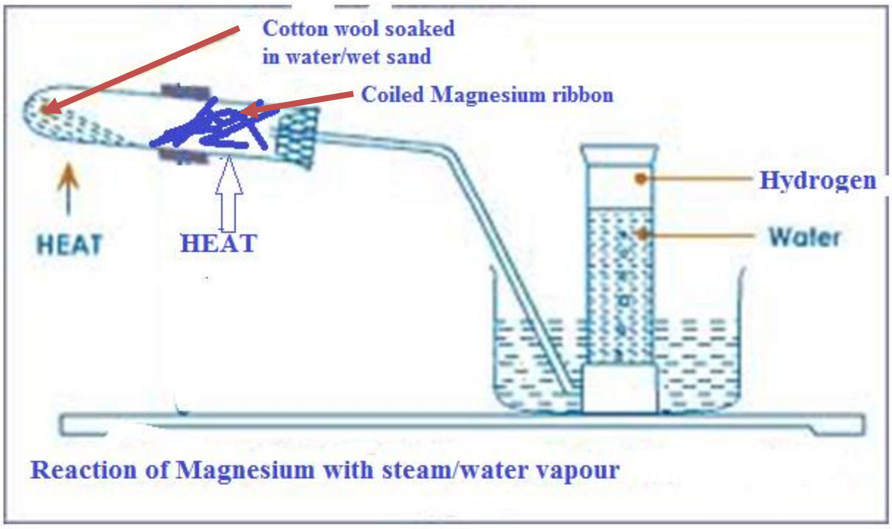
$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

(iv)Reaction with water vapour/steam

Experiment

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

Set up of apparatus



Reaction of Magnesium with steam/water vapour

Observations

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

Explanation

On heating wet sand, steam is generated which drives out the air that would otherwise react with /oxidize the ribbon.

Magnesium burns in steam/water vapour generating enough heat that ensures the reaction goes to completion even without further heating. White Magnesium oxide is formed and hydrogen gas is evolved.

To prevent suck back, the delivery tube should be removed from the water before heating is stopped at the end of the experiment.

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

(v)Reaction with chlorine gas.

Experiment

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

Observation

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

Explanation

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

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

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

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

(v)Reaction with dilute acids.

Experiment

Place about 4.0cm3 of 0.1M dilute sulphuric(VI)acid into a test tube. Add about 1.0cm length of magnesium ribbon into the test tube. Cover the mouth of the test tube using a thumb. Release the gas and test the gas using a burning splint.

Repeat with about 4.0cm3 of 0.1M dilute hydrochloric/nitric(V) acid.

Repeat with 0.1g of Calcium in a beaker with all the above acid

Caution: Keep distance when using calcium

Observation

-Effervescence/fizzing/bubbles with dilute sulphuric(VI) and nitric(V) acids

-Little Effervescence/fizzing/bubbles with calcium and dilute sulphuric(VI) acid.

-Colourless gas produced that extinguishes a burning splint with an explosion/ “pop” sound.

-No gas is produced with Nitric(V)acid.

-Colourless solution is formed.

Explanation

Dilute acids react with alkaline earth metals to form a salt and produce hydrogen gas.

Nitric(V)acid is a strong oxidizing agent. It quickly oxidizes the hydrogen produced to water.

Calcium is very reactive with dilute acids and thus a very **small** piece of very **dilute** acid should be used.

Chemical equations

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

$$Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(\mathbf{s}) + H_2(g)$$

(insoluble $CaSO_4$(**s**) coat/cover Ca(s))

$$Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$$
$$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$$

$$Ba(s) + H_2SO_4(aq) \rightarrow BaSO_4(\mathbf{s}) + H_2(g)$$

(insoluble $BaSO_4$(**s**) coat/cover Ba(s))

$$Ba(s) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + H_2(g)$$
$$Ba(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2(g)$$

The table below shows some compounds of some alkaline earth metals

	Beryllium	Magnesium	Calcium	Barium
Hydroxide	$Be(OH)_2$	$Mg(OH)_2$	$Ca(OH)_2$	$Ba(OH)_2$
Oxide	BeO	MgO	CaO	BaO

Sulphide	-	MgS	CaS	BaS
Chloride	$BeCl_2$	$MgCl_2$	$CaCl_2$	$BaCl_2$
Carbonate	$BeCO_3$	$MgCO_3$	$CaCO_3$	$BaCO_3$
Nitrate(V)	$Be(NO_3)_2$	$Mg(NO_3)_2$	$Ca(NO_3)_2$	$Ba(NO_3)_2$
Sulphate(VI)	$BeSO_4$	$MgSO_4$	$CaSO_4$	$BaSO_4$
Sulphate(IV)	-	-	$CaSO_3$	$BaSO_3$
Hydrogen carbonate	-	$Mg(HCO_3)_2$	$Ca(HCO_3)_2$	-
Hydrogen sulphate(VI)	-	$Mg(HSO_4)_2$	$Ca(HSO_4)_2$	-

Some uses of alkaline earth metals include:

(i)Magnesium hydroxide is a non-toxic/poisonous mild base used as an anti acid medicine to relieve stomach acidity.

(ii)Making **duralumin.** Duralumin is an alloy of Magnesium and aluminium used for making aeroplane bodies because it is light.

(iii) Making plaster of Paris-Calcium sulphate(VI) is used in hospitals to set a fractures bone.

(iii)Making cement-Calcium carbonate is mixed with clay and sand then heated to form cement for construction/building.

(iv)Raise soil pH-Quicklime/calcium oxide is added to acidic soils to neutralize and raise the soil pH in agricultural farms.

(v)As nitrogenous fertilizer-Calcium nitrate(V) is used as an agricultural fertilizer because plants require calcium for proper growth.

(vi)In the blast furnace-Limestone is added to the blast furnace to produce more reducing agent and remove slag in the blast furnace for extraction of Iron.

(c)Group VII elements: Halogens

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

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

All halogen atoms have seven electrons in the outer energy level. They **acquire/gain one** electron in the outer energy level to be stable. They therefore are therefore **monovalent** .They exist in oxidation state X^-

The number of energy levels increases down the group from Fluorine to Astatine. The more the number of energy levels the bigger/larger the atomic size. e.g.

The atomic size/radius of Chlorine is bigger/larger than that of Fluorine because Chlorine has more/3 energy levels than Fluorine (2 energy levels).

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

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

Table showing the atomic and ionic radius of four Halogens

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

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

Electronegativity

The ease of gaining/acquiring extra electrons is called electronegativity. All halogens are electronegative. Electronegativity decreases as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius.

The outer electrons experience less nuclear attraction and thus ease of gaining/acquiring extra electrons decrease.

It is measured using Pauling's scale.

Where Fluorine with Pauling scale 4.0 is the most electronegative element and thus the highest tendency to acquire/gain extra electron.

Table showing the electronegativity of the halogens.

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

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

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

Electron affinity

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

Table showing the election affinity of halogens for the process

$X + e \rightarrow X^-$

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

The higher the electron affinity the more stable theion.i.e

Cl^- is a more stable ion than Br^- because it has a more negative / exothermic electron affinity than Br^-

Electron affinity is different from:

(i) Ionization energy.

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

(ii) Electronegativity.

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

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

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

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

Physical properties

State at room temperature

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

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

Melting/Boiling point

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

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

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

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

Electrical conductivity

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

Solubility in polar and non-polar solvents

All halogens are soluble in water(polar solvent).

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

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

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

Caution: Tetrachloromethane , Bromine vapour and Chlorine gas are all **highly** toxic/poisonous.

Table showing the physical properties of Halogens

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

Chemical properties

(i)Displacement

Experiment

Place separately in test tubes about $5cm^3$ of sodium chloride, Sodium bromide and Sodium iodide solutions.

Add 5 drops of chlorine water to each test tube:

Repeat with 5 drops of bromine water instead of chlorine water

Observation

Using Chlorine water

-Yellow colour of chlorine water fades in all test tubes except with sodium chloride.

-Coloured Solution formed.

Using Bromine water

Yellow colour of bromine water fades in test tubes containing sodium iodide.

-Coloured Solution formed.

Explanation

The halogens displace each other from their solution. The more electronegative displace the less electronegative from their solution.

Chlorine is more electronegative than bromine and iodine.

On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.

Bromine is more electronegative than iodide but less 6than chlorine.

On adding Bromine water, iodine is displaced from its solution but not chlorine.

Table showing the displacement of the halogens

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

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

Chemical /ionic equations

With Fluorine

$F_2(g)$ + $2NaCl^-(aq)$ -> $2NaF(aq)$ + $Cl_2(aq)$
$F_2(g)$ + $2Cl^-(aq)$ -> $2F^-(aq)$ + $Cl_2(aq)$

$F_2(g)$ + $2NaBr^-(aq)$ -> $2NaF(aq)$ + $Br_2(aq)$
$F_2(g)$ + $2Br^-(aq)$ -> $2F^-(aq)$ + $Br_2(aq)$

$F_2(g)$ + $2NaI^-(aq)$ -> $2NaF(aq)$ + $I_2(aq)$
$F_2(g)$ + $2I^-(aq)$ -> $2F^-(aq)$ + $I_2(aq)$

With chlorine

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

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

With Bromine

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

Uses of halogens

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

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

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

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

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

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

The table below to show some compounds of halogens.

Element / Halogen	H	Na	Mg	Al	Si	C	P
F	**HF**	**NaF**	**MgH_2**	**AlF_3**	**SiF_4**	**CF_4**	**PF_3**
Cl	**HCl**	**$NaCl$**	**$MgCl$**	**$AlCl_3$**	**$SiCl_3$**	**CCl_4**	**PCl_3**
Br	**HBr**	**$NaBr$**	**$MgBr_2$**	**$AlBr_3$**	**$SiBr_4$**	**CBr_4**	**PBr_3**
I	**Hl**	**Nal**	**Mgl_2**	**All_3**	**SiI_4**	**Cl_2**	**PBr_3**

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

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

I. What do you understand by the term "bond energy"

Bond energy is the energy required to break/ form one mole of chemical bond

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

(c)Group VIII elements: Noble gases

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

Element	Symbol	Atomic number	Electron structure	State at room temperature
Helium	He	2	2:	Colourless gas
Neon	Ne	10	2:8	Colourless gas
Argon	Ar	18	2:8:8	Colourless gas
Krypton	Kr	36	2:8:18:8	Colourless gas
Xenon	Xe	54	2:8:18:18:8	Colourless gas
Radon	Rn	86	2:8:18:32:18:8	Radioctive

All noble gas atoms have a stable duplet(two electrons in the 1st energy level) or octet(eight electrons in other outer energy level)in the outer energy level. They therefore **do not** acquire/gain extra electron in the outer energy level or donate/lose. They therefore are therefore **zerovalent** .
The number of energy levels increases down the group from Helium to Randon. The more the number of energy levels the bigger/larger the atomic size/radius. e.g.
The atomic size/radius of Argon is bigger/larger than that of Neon because Argon has more/3 energy levels than Neon (2 energy levels).

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

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

Noble gases have low melting and boiling points. This is because they exist as monatomic molecules joined by very weak intermolecular/van-der-waals forces that require very little energy to weaken and form liquid and break to form a gas.

The intermolecular/van-der-waals forces increase down the group as the atomic radius/size increase from Helium to Radon. The melting and boiling points thus increase also down the group.

Noble gases are insoluble in water and are poor conductors of electricity.

Element	Formula of molecule	Electrical conductivity	Solubility in water	Atomic radius(nM)	1st ionization energy	Melting point(^{0}C)	Boiling point(^{0}C)
Helium	He	Poor	Insoluble	0.128	2372	-270	-269
Neon	Ne	Poor	Insoluble	0.160	2080	-249	-246
Argon	Ar	Poor	Insoluble	0.192	1520	-189	-186
Krypton	Kr	Poor	Insoluble	0.197	1350	-157	-152
Xenon	Xe	Poor	Insoluble	0.217	1170	-112	-108
Radon	Rn	Poor	Insoluble	0.221	1134	-104	-93

Uses of noble gases

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

Argon is used in arch welding as an insulator.

Neon is used in street and advertisement light

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

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

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

C. PERIODICITY OF ACROSS THE PERIOD.

(See Chemical bonding and Structure)

UPGRADE
CHEMISTRY
FORM 2

Periodicity of CHEMICAL FAMILIES

Comprehensive tutorial notes

MUTHOMI S.G
www.kcselibrary.info
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PERIODICITY OF CHEMICAL FAMILES
(Patterns down the group)

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

(b) Group I elements: Alkali metals

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

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

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

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

The distance between the centre of the nucleus of an **ion** and the outermost energy level occupied by electron/s is called **ionic radius.** Ionic radius is also measured in **nanometers**(n).The higher /bigger the ionic radius the bigger /larger the size of the ion.
Atomic radius and ionic radius depend on the number of energy levels occupied by electrons. The more the number of energy levels the bigger/larger the atomic /ionic radius. e.g.
The atomic radius of Francium is bigger/larger than that of sodium because Francium has more/7 energy levels than sodium (3 energy levels).

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

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

Table showing the atomic and ionic radius of some alkali metals

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

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

Electropositivity

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

Ionization energy

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

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

Physical properties

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

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

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

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

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

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

Chemical properties

(i)Reaction with air/oxygen

On exposure to air, alkali metals reacts with the elements in the air.

Example

On exposure to air, Sodium first reacts with Oxygen to form sodium oxide.

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

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

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

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

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

(ii)Burning in air/oxygen

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

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

Sodium burns in air with a **yellow** flame to form sodium oxide

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

Sodium burns in oxygen with a **yellow** flame to form sodium peroxide

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

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

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

(iii) Reaction with water:

Experiment

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

Observations

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

Explanation

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

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

Chemical equations

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

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

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

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

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

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

Reactivity increase down the group

(iv) Reaction with chlorine:

Experiment

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

Repeat with about 0.5g of Lithium.

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

Observation

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

Lithium metal continues to burn with a crimson flame forming white solid / fumes.

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

Chemical equations

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

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

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

$2Rb(s) + Cl_2(g) \rightarrow 2RbCl(s)$

$2Cs(s) + Cl_2(g) \rightarrow 2CsCl(s)$

$2Fr(s) + Cl_2(g) \rightarrow 2FrCl(s)$

Reactivity increase down the group

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

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

Some **uses** of alkali metals include:
(i)Sodium is used in making sodium cyanide for extracting gold from gold ore.
(ii)Sodium chloride is used in seasoning food.
(iii)Molten mixture of sodium and potassium is used as coolant in nuclear reactors.
(iv)Sodium is used in making sodium hydroxide used in making soapy and soapless detergents.
(v)Sodium is used as a reducing agent for the extraction of titanium from Titanium(IV)chloride.
(vi)Lithium is used in making special high strength glasses
(vii)Lithium compounds are used to make dry cells in mobile phones and computer laptops.

Group II elements: Alkaline earth metals

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

Element	Symbol	Atomic number	Electron structure	Oxidation state	Valency
Beryllium	Be	4	2:**2**	Be^{2+}	2
Magnesium	Mg	12	2:8:**2**	Mg^{2+}	2
Calcium	Ca	20	2:8:8:**2**	Ca^{2+}	2

Strontium	Sr	38	2:8:18:8:**2**	Sr^{2+}	2
Barium	Ba	56	2:8:18:18:8:**2**	Ba^{2+}	2
Radium	Ra	88	2:8:18:32:18:8:**2**	Ra^{2+}	2

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

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

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

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

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

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

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

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

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

Physical properties

Soft/Easy to cut: Alkaline earth metals are **not** soft and easy to cut with a knife like alkali metals. This is because of the decrease in atomic radius of corresponding alkaline earth metal, increases the strength of metallic bond and the packing of the metallic structure. Alkaline earth metals are
(i)ductile(able to form wire/thin long rods)
(ii)malleable(able to be hammered into sheet/long thin plates)
(iii)have high tensile strength(able to be coiled without breaking/ not brittle/withstand stress)

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

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

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

Electrical/thermal conductivity: Alkaline earth metals are good thermal and electrical conductors. The two delocalized valence electrons move randomly within the metallic structure.

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

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

Calcium is a better conductor than magnesium.

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

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

Alkaline earth metal	Appearance	Ease of cutting	Melting point (°C)	Boiling point (°C)	Conduct - ivity	1st ionization energy	2nd ionization energy
Beryllium	Shiny grey	Not easy	1280	3450	Good	900	1800
Magnesium	Shiny grey	Not Easy	650	1110	Good	736	1450
calcium	Shiny grey	Not easy	850	1140	Good	590	970

Chemical properties

(i)Reaction with air/oxygen

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

Example

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

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

(ii)Burning in air/oxygen

Experiment

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

Caution: Do not look directly at the flame

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

Repeat with Calcium

Observations

-Magnesium burns with a bright blindening flame
-White solid /ash produced
-Solid dissolves in water to form a colourless solution
-Blue litmus paper remain blue

-Red litmus paper turns blue
-colourless gas with pungent smell of urine

Explanation

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

$2Mg (s) + O_2(g) \rightarrow 2MgO(s)$
$3Mg (s) + N_2 (g) \rightarrow Mg_3N_2 (s)$

Magnesium oxide dissolves in water to form magnesium hydroxide.

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

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

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

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

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

$2Ca (s) + O_2(g) \rightarrow 2CaO(s)$
$3Ca (s) + N_2 (g) \rightarrow Ca_3N_2 (s)$

Calcium oxide dissolves in water to form calcium hydroxide.

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

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

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

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

(iii)Reaction with water

Experiment

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

Observations

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

Explanations

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

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

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

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

(iv)Reaction with water vapour/steam

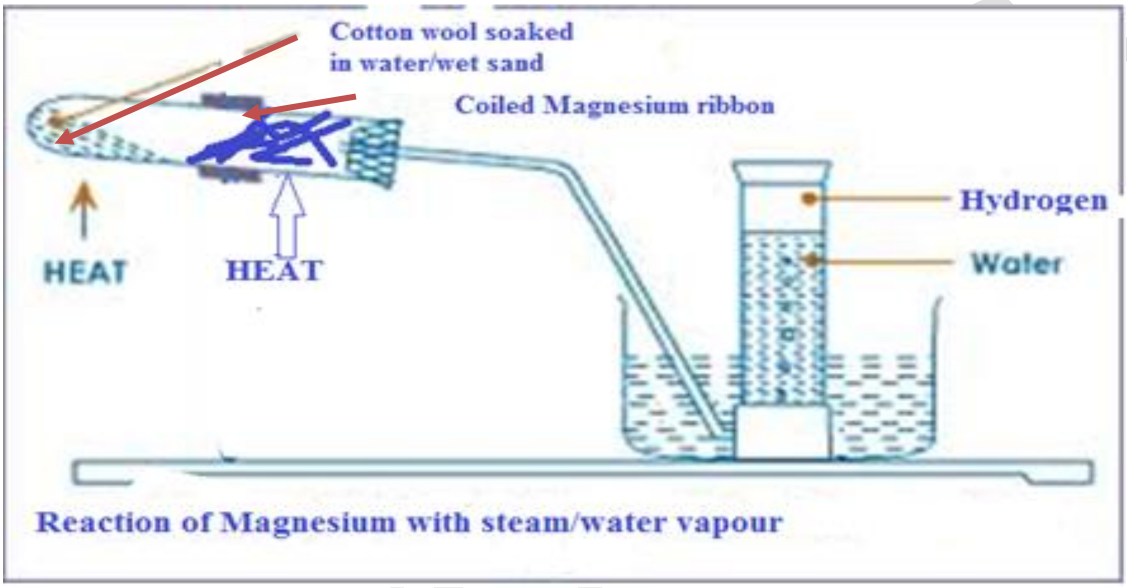
Experiment

Put some cotton wool soaked in water/wet sand in a long boiling tube.

Coil a well polished magnesium ribbon into the boiling tube.

Ensure the coil touches the side of the boiling tube. Heat the cotton wool/sand slightly then strongly heat the Magnesium ribbon .

Set up of apparatus



Reaction of Magnesium with steam/water vapour

Observations

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

Explanation

On heating wet sand, steam is generated which drives out the air that would otherwise react with /oxidize the ribbon.

Magnesium burns in steam/water vapour generating enough heat that ensures the reaction goes to completion even without further heating. White Magnesium oxide is formed and hydrogen gas is evolved.

To prevent suck back, the delivery tube should be removed from the water before heating is stopped at the end of the experiment.

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

(v)Reaction with chlorine gas.

Experiment

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

Observation

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

Explanation

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

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

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

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

(v)Reaction with dilute acids.

Experiment

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

Observation

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

Explanation

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

Chemical equations

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

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

$$Ba(s) + H_2SO_4(aq) \rightarrow BaSO_4(\mathbf{s}) + H_2(g)$$

(insoluble $BaSO_4$(**s**) coat/cover Ba(s))

$$Ba(s) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + H_2(g)$$

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

The table below shows some compounds of some alkaline earth metals

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

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

(c)Group VII elements: Halogens

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

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

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

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

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

Table showing the atomic and ionic radius of four Halogens

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

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

Electronegativity

The ease of gaining/acquiring extra electrons is called electronegativity. All halogens are electronegative. Electronegativity decreases as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius.
The outer electrons experience less nuclear attraction and thus ease of gaining/acquiring extra electrons decrease.

It is measured using Pauling's scale.

Where Fluorine with Pauling scale 4.0 is the most electronegative element and thus the highest tendency to acquire/gain extra electron.

Table showing the electronegativity of the halogens.

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

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

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

Electron affinity

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

Table showing the election affinity of halogens for the process

$$X + e \rightarrow X^-$$

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

The higher the electron affinity the more stable theion.i.e

Cl^- is a more stable ion than Br^- because it has a more negative / exothermic electron affinity than Br^-

Electron affinity is different from:

(i) Ionization energy.

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

(ii) Electronegativity.

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

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

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

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

Physical properties

State at room temperature

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

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

Melting/Boiling point

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

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

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

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

Electrical conductivity

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

Solubility in polar and non-polar solvents

All halogens are soluble in water(polar solvent).

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

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

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

Caution: Tetrachloromethane , Bromine vapour and Chlorine gas are all **highly** toxic/poisonous.

Table showing the physical properties of Halogens

Halogen	**Formula of molecule**	**Electrical conductivity**	**Solubility in water**	**Melting point(°C)**	**Boiling point(°C)**
Fluorine	F_2	Poor	Insoluble/soluble in tetrachloromethane	-238	-188
Chlorine	Cl_2	Poor	Insoluble/soluble	-101	-35

			in tetrachloromethane		
Bromine	Br_2	Poor	Insoluble/soluble in tetrachloromethane	7	59
Iodine	I_2	Poor	Insoluble/soluble in tetrachloromethane	114	sublimes

Chemical properties

(i)Displacement

Experiment

Place separately in test tubes about 5cm^3 of sodium chloride, Sodium bromide and Sodium iodide solutions.

Add 5 drops of chlorine water to each test tube:

Repeat with 5 drops of bromine water instead of chlorine water

Observation

Using Chlorine water

-Yellow colour of chlorine water fades in all test tubes except with sodium chloride.

-Coloured Solution formed.

Using Bromine water

Yellow colour of bromine water fades in test tubes containing sodium iodide.

-Coloured Solution formed.

Explanation

The halogens displace each other from their solution. The more electronegative displace the less electronegative from their solution.

Chlorine is more electronegative than bromine and iodine.

On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.

Bromine is more electronegative than iodide but less 6than chlorine.

On adding Bromine water, iodine is displaced from its solution but not chlorine.

Table showing the displacement of the halogens

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

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

Chemical /ionic equations

With Fluorine

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

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

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

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

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

With Bromine
$Br_2(g) + 2NaI^-(aq) \rightarrow 2NaBr(aq) + I_2(aq)$
$Br_2(g) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$

Uses of halogens

(iv) Florine – manufacture of P.T.F.E (Poly tetra fluoroethene) synthetic fiber.
- Reduce tooth decay when added in small amounts/quantities in tooth paste.

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

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

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

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

The table below to show some compounds of halogens.

Element / Halogen	H	Na	Mg	Al	Si	C	P
F	**HF**	**NaF**	**MgF_2**	**AlF_3**	**SiF_4**	**CF_4**	**PF_3**
Cl	**HCl**	**NaCl**	**$MgCl_2$**	**$AlCl_3$**	**$SiCl_4$**	**CCl_4**	**PCl_3**

Br	**HBr**	**NaBr**	**$MgBr_2$**	**$AlBr_3$**	**$SiBr_4$**	**CBr_4**	**PBr_3**
I	**Hl**	**Nal**	**Mgl_2**	**All_3**	**SiI_4**	**$C\ l_4$**	**PBr_3**

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

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

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

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

(c)Group VIII elements: Noble gases

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

Element	Symbol	Atomic number	Electron structure	State at room temperature
Helium	He	2	2:	Colourless gas
Neon	Ne	10	2:8	Colourless gas
Argon	Ar	18	2:8:8	Colourless gas
Krypton	Kr	36	2:8:18:8	Colourless gas
Xenon	Xe	54	2:8:18:18:8	Colourless gas
Radon	Rn	86	2:8:18:32:18:8	Radioctive

All noble gas atoms have a stable duplet(two electrons in the 1st energy level) or octet(eight electrons in other outer energy level)in the outer energy level. They therefore **do not** acquire/gain extra electron in the outer energy level or donate/lose. They therefore are therefore **zerovalent** .
The number of energy levels increases down the group from Helium to Randon. The more the number of energy levels the bigger/larger the atomic size/radius. e.g.
The atomic size/radius of Argon is bigger/larger than that of Neon because Argon has more/3 energy levels than Neon (2 energy levels).

Atomic radius noble gases increase down the group as the number of energy levels increases.

The effective nuclear attraction on the outer electrons thus decrease down the group.

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

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

Element	Formula of molecule	Electrical conductivity	Solubility in water	Atomic radius(nM)	1st ionization energy	Melting point(^{0}C)	Boiling point(^{0}C)
Helium	He	Poor	Insoluble	0.128	2372	-270	-269
Neon	Ne	Poor	Insoluble	0.160	2080	-249	-246
Argon	Ar	Poor	Insoluble	0.192	1520	-189	-186
Krypton	Kr	Poor	Insoluble	0.197	1350	-157	-152
Xenon	Xe	Poor	Insoluble	0.217	1170	-112	-108
Radon	Rn	Poor	Insoluble	0.221	1134	-104	-93

Uses of noble gases
Argon is used in light bulbs to provide an inert environment to prevent oxidation of the bulb filament
Argon is used in arch welding as an insulator.
Neon is used in street and advertisement light
Helium is mixed with Oxygen during deep sea diving and mountaineering.
Helium is used in weather balloon for meteorological research instead of Hydrogen because it is unreactive/inert. Hydrogen when impure can ignite with an explosion.
Helium is used in making thermometers for measuring very low temperatures.

UPGRADE CHEMISTRY

FORM 2

STRUCTURE & BONDING

Comprehensive tutorial notes

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CHEMICAL BONDING AND STRUCTURE

A. CHEMICAL BONDING

A chemical bond is formed when atoms of the same or different elements **share**, **gain**, **donate or delocalize** their **outer** energy level **electrons** to combine during **chemical reactions** inorder to be **stable**.

Atoms have equal number of negatively charged electrons in the energy levels and positively charged protons in the nucleus.

Atoms are chemically stable if they have filled outer energy level. An energy level is full if it has duplet (2) or octet (8) state in outer energy level.

Noble gases have duplet /octet. All other atoms try to be like noble gases through chemical reactions and forming molecules.

Only electrons in the outer energy level take part in formation of a chemical bond. There are three main types of chemical bonds formed by atoms:

(i) covalent bond
(ii) ionic/electrovalent bond
(iii) metallic bond

(i)COVALENT BOND

A covalent bond is formed when atoms of the same or different element share some or all the outer energy level electrons to combine during chemical reactions inorder to attain duplet or octet.

A shared pair of electrons is attracted by the nucleus (protons) of the two atoms sharing.

Covalent bonds are mainly formed by non-metals to form molecules. A molecule is a group of atoms of the same or different elements held together by a covalent bond. The number of atoms making a molecule is called **atomicity**. Noble gases are **monatomic** because they are stable and thus do not bond with each other or other atoms. Most other gases are **diatomic**

The more the number of electrons shared, the stronger the covalent bond.

A pair of electrons that do not take part in the formation of a covalent bond is called a **lone pair of electrons.**

Mathematically, the number of electrons to be shared by an atom is equal to the number of electrons remaining for the atom to be stable/attain duplet/octet /have maximum electrons in outer energy level.

The following diagrams illustrate the formation of covalent bonds:

a)hydrogen molecule is made up of two hydrogen atoms in the outer energy level each requiring one electron to have a stable duplet.

To show the formation of covalent bonding in the molecule then the following data/information is required;

Symbol of atom/element taking part in bonding	H	H
Number of protons/electrons	1	1
Electron configuration/structure	1:	1:
Number of electron in outer energy level	1	1

Number of electrons remaining to be stable/shared	1	1
Number of electrons not shared(lone pairs)	0	0

Diagram method 1

Diagram method 2

H ●x

Note:

After bonding the following **intramolecular** forces exist:

(i)the attraction of the shared electrons by both nucleus /protons of the atoms

(ii) the repulsion of the nucleus of one atom on the other.

(iii)balance of the attraction and repulsion is maintained inside/intramolecular/within the molecule as follows;

E_1

P_1 P_1

E_1

(iv)Protons(P_1) from nucleus of atom 1 **repel** protons (P_2) from nucleus of atom 2.

(v)Electron (E_1) in the energy levels of atom 1 **repel** electron (E_2) in the energy levels of atom 2.

(vi) Protons(P_1) from nucleus of atom 1 **attract** electron (E_2) in the energy levels of atom 2.

(vii) protons (P_2) from nucleus of atom 2 **attract** electron (E_2) in the energy levels of atom 2.

b) Fluorine, chlorine, bromine and iodine molecules are made up also of two atoms sharing the outer energy level electrons to have a stable octet.

To show the formation of covalent bonding in the molecule then the following data/information is required;

(i) fluorine

Symbol of atom/element taking part in bonding	F	F
Number of protons/electrons	9	9
Electron configuration/structure	2:7	2:7

Number of electron in outer energy level	7	7
Number of electrons remaining to be stable/shared	1	1
Number of outer electrons not shared(3-lone pairs)	6	6

Diagram method 1

Diagram method 2

(ii) chlorine

Symbol of atom/element taking part in bonding	Cl	Cl
Number of protons/electrons	17	17
Electron configuration/structure	2:8:7	2:8:7
Number of electron in outer energy level	7	7
Number of electrons remaining to be stable/shared	1	1
Number of outer electrons not shared(3-lone pairs)	6	6

Diagram method 1

Diagram method 2

(iii) Bromine

Symbol of atom/element taking part in bonding	Br	Br
Number of protons/electrons	35	35
Electron configuration/structure	2:8:18:7	2:8:18:7
Number of electron in outer energy level	7	7

Number of electrons remaining to be stable/shared	1	1
Number of outer electrons not shared(3-lone pairs)	6	6

Diagram method 1

Diagram method 2

(iv) Iodine

Symbol of atom/element taking part in bonding	I	I
Number of protons/electrons	53	53
Electron configuration/structure	2:8:18:18:7	2:8:18:18:7
Number of electron in outer energy level	7	7
Number of electrons remaining to be stable/shared	1	1
Number of outer electrons not shared(3-lone pairs)	6	6

Diagram method 1

Diagram method 2

c) Oxygen molecule is made up of two atoms sharing each two outer energy level electrons to have a stable octet as shown below;

Symbol of atom/element taking part in bonding	O	O
Number of protons/electrons	8	8
Electron configuration/structure	2:6	2:6
Number of electron in outer energy level	6	6
Number of electrons remaining to be stable/shared	2	2
Number of outer electrons not shared(2-lone pairs)	4	4

Diagram method 1

Diagram method 2

d) Nitrogen and phosphorus molecule is made up of two atoms sharing each three outer energy level electrons to have a stable octet as shown below;

(i) Nitrogen

Symbol of atom/element taking part in bonding	N	N
Number of protons/electrons	7	7
Electron configuration/structure	2:5	2:5
Number of electron in outer energy level	5	5
Number of electrons remaining to be stable/shared	3	3
Number of outer electrons not shared (3-lone pairs)	2	2

Diagram method 1

Diagram method 2

(ii) Phosphorus

Symbol of atom/element taking part in bonding	P	P
Number of protons/electrons	15	15
Electron configuration/structure	2:8:5	2:8:5
Number of electron in outer energy level	5	5
Number of electrons remaining to be stable/shared	3	3
Number of outer electrons not shared (3-lone pairs)	2	2

Diagram method 1

Diagram method 2

e) Water molecule is made up of hydrogen and oxygen. Hydrogen requires to share one electron with oxygen to be stable/attain duplet. Oxygen requires to share two electrons to be stable/attain octet. Two hydrogen atoms share with one oxygen atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	O	H
Number of protons/electrons	8	1
Electron configuration/structure	2:6	1
Number of electron in outer energy level	6	1
Number of electrons remaining to be stable/shared	2	1
Number of electrons not shared(2-Oxygen lone pairs)	4	0

Diagram method 1

Diagram method 2

f) Ammonia molecule is made up of Hydrogen and Nitrogen. Hydrogen requires to share one electron with Nitrogen to be stable/attain duplet. Nitrogen requires to share three electrons to be stable/attain octet. Three hydrogen atoms share with one nitrogen atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	N	H
Number of protons/electrons	7	1
Electron configuration/structure	2:5	1:
Number of electron in outer energy level	5	1
Number of electrons remaining to be stable/shared	3	1
Number of electrons not shared(1-Nitrogen lone pairs)	2	0

Diagram method 1

Diagram method 2

g)Carbon(IV) oxide molecule is made up of carbon and oxygen. Carbon requires to share four electrons with oxygen to be stable/attain octet. Oxygen requires to share two electrons to be stable/attain octet. Two oxygen atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	O	C
Number of protons/electrons	8	6
Electron configuration/structure	2:6	2:4
Number of electron in outer energy level	6	4
Number of electrons remaining to be stable/shared	2	4
2-lone pairs from each Oxygen atom)	2	0

Diagram method 1

Diagram method 2

h) Methane molecule is made up of hydrogen and carbon. Hydrogen requires sharing one electron with carbon to be stable/attain duplet. Carbon requires sharing four electrons to be stable/attain octet. Four hydrogen atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	C	H
Number of protons/electrons	6	1
Electron configuration/structure	2:4	1
Number of electron in outer energy level	4	1

Number of electrons remaining to be stable/shared	4	1
Number of electrons not shared (No lone pairs)	0	0

Diagram method 1

Diagram method 2

i) Tetrachloromethane molecule is made up of chlorine and carbon. Chlorine requires sharing one electron with carbon to be stable/attain octet. Carbon requires sharing four electrons to be stable/attain octet. Four chlorine atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	C	Cl
Number of protons/electrons	6	17
Electron configuration/structure	2:4	2:8:7
Number of electron in outer energy level	4	7
Number of electrons remaining to be stable/shared	4	1
3-lone pairs from each Chlorine atom(24 electrons)	0	6

Diagram method 1

Diagram method 2

j) Ethane molecule is made up of six hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Three hydrogen atoms share with one carbon atom while another three hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing a pair of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding	C	H
Number of protons/electrons	6	1
Electron configuration/structure	2:4	1
Number of electron in outer energy level	4	1
Number of electrons remaining to be stable/shared	4	1
Number of electrons not shared(No lone pairs)	0	0

Diagram method 1

Diagram method 2

k) Ethene molecule is made up of four hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Two hydrogen atoms share with one carbon atom while another two hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing two pairs of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding	C	H
Number of protons/electrons	6	1
Electron configuration/structure	2:4	1
Number of electron in outer energy level	4	1
Number of electrons remaining to be stable/shared	4	1
Number of electrons not shared(No lone pairs)	0	0

Diagram method 1

Diagram method 2

l) Ethyne molecule is made up of two hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. One hydrogen atoms share with one carbon atom while another hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing three pairs of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding	C	H
Number of protons/electrons	6	1
Electron configuration/structure	2:4	1
Number of electron in outer energy level	4	1
Number of electrons remaining to be stable/shared	4	1
Number of electrons not shared(No lone pairs)	0	0

Diagram method 1

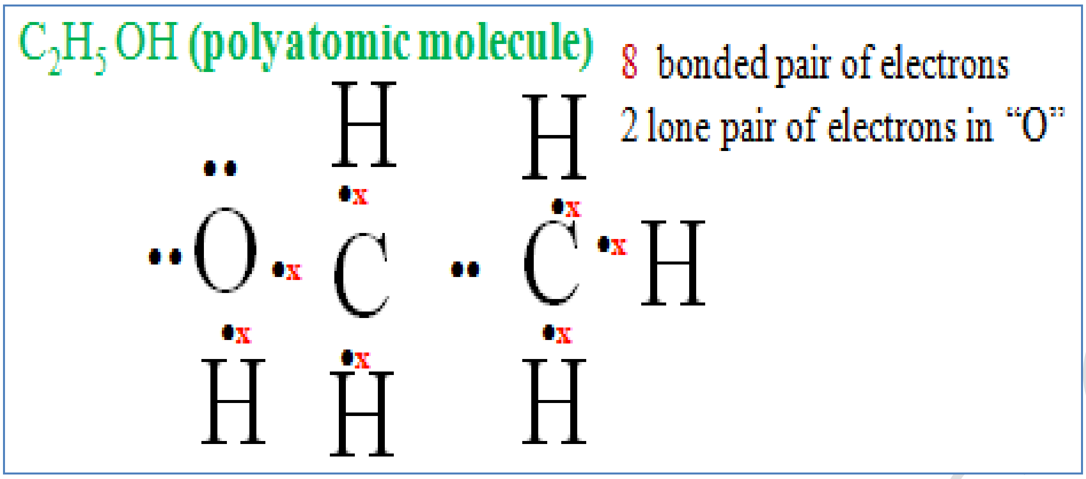
Diagram method 2

j) Ethanol molecule is made up of six hydrogen one Oxygen atom two carbon atoms.

Five Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet

Each Carbon uses four electrons to share with “O”and “H”attain octet/duplet.

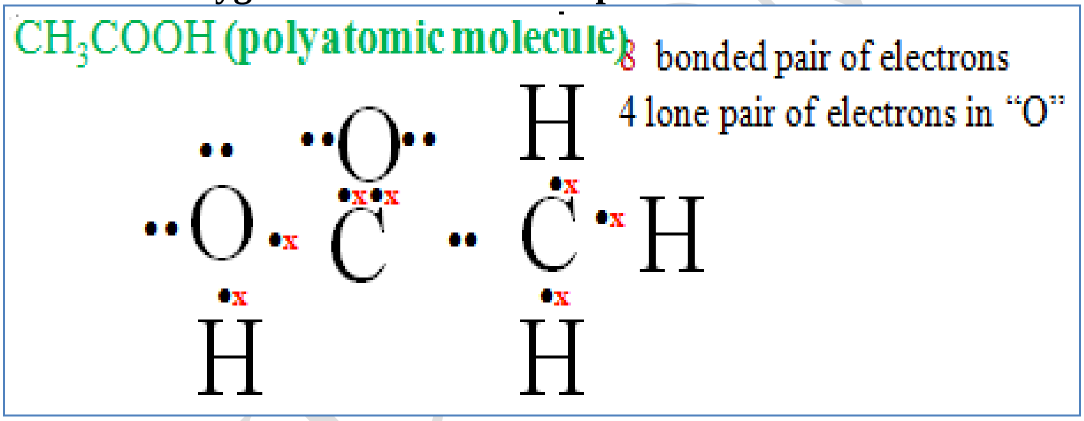
NB: Oxygen has two lone pairs



j)Ethanoic molecule is made up of four hydrogen two Oxygen atom two carbon atoms. Three Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet

Each Carbon uses four electrons to share with "O"and "H"attain octet/duplet.

NB: Each Oxygen atom has two lone pairs



By convention (as a rule), a

(i) **single** covalent bond made up of two shared(**a pair**) electrons is represented by a dash(---)

(ii) **double** covalent bond made up of four shared(**two pairs)** electrons is represented by a double dash(==)

(iii) **triple** covalent bond made up of six shared(**three pairs)** electrons is represented by a triple dash(==)

The representation below show the molecules covered in (a) to (k) above:

a) Hydrogen molecule(H_2) H--H

b) Fluorine molecule(F_2)	F--F
c) Chlorine molecule(Cl_2)	Cl--Cl
d) Bromine molecule(Br_2)	Br--Br
e) Iodine molecule(I_2)	I--I
f) Oxygen molecule(O_2)	O=O
g) Nitrogen molecule(N_2)	N=N
h) Phosphorus molecule(P_2)	P=P
i) Water molecule (H_2O)	H--O--H
j Ammonia molecule(NH_3)	H--N--H H
k)Carbon(IV) oxide molecule(CO_2)	O==C==O
l)Methane molecule(CH_4)	H H--C--H H
m)Tetrachloromethane molecule(CCl_4	Cl Cl--C--Cl Cl
n)Ethane molecule(C_2H_6)	H H H--C—C--H H H
p)Ethene molecule(C_2H_4)	H-C==C-H H H
q)Ethyne molecule(C_2H_6)	H-C—C-H

Dative /coordinate bond

A dative/coordinate bond is a covalent bond formed when a lone pair of electrons is donated then shared to an electron-deficient species/ion/atom.

During dative/coordinate bonding, all the shared pair of electrons are donated by one of the combining/bonding species/ ion/atom.

Like covalent bonding, coordinate /dative bond is mainly formed by non-metals.

Illustration of coordinate /dative bond

a)Ammonium ion(NH_4^+)

The ammonium ion is made up of ammonia (NH_3) molecule and hydrogen (H^+) ion. (H^+) ion has no electrons. NH_3 is made up of covalent bonding from Nitrogen and Hydrogen. One lone pair of electrons is present in Nitrogen atom after the bonding. This lone pair is donated and shared with an electron-deficient H^+ ion

Diagram method 1

Diagram method 2

b)Phosphine ion (PH_4^+)

The Phosphine ion is made up of phosphine(NH_3) molecule and hydrogen (H^+) ion. (H^+) ion has no electrons. PH_3 is made up of covalent bonding from Phosphorus and Hydrogen. One lone pair of electrons is present in Phosphorus atom. After the bonding this lone pair is donated and shared with the electron-deficient H^+ ion

Diagram method 1

Diagram method 2

c) Hydroxonium (H_3O^+) ion

The hydroxonium ion is made up of water (H_2O) molecule and hydrogen (H^+) ion. (H^+) ion has no electrons. The H_2O molecule is made up of covalent bonding from Oxygen and Hydrogen. One lone pair of electrons out of the two present in Oxygen atom after the bonding is donated and shared with the electron-deficient H^+ ion

Diagram method 1

Diagram method 2

d) Carbon (II) oxide (CO)

Carbon (II) oxide is made up of carbon and Oxygen atoms sharing each two outer electron and not sharing each two electrons. Oxygen with an extra lone pair of electrons donates and share with the carbon atom for both to be stable.

Diagram method 1

Diagram method 2

e) Aluminium (III) chloride ($AlCl_3/Al_2Cl_6$)

Aluminium (III) chloride is made up of aluminium and chlorine. One aluminium atom shares its outer electrons with three separate chlorine atoms. All chlorine atoms attain stable octet but aluminium does not. Another molecule of aluminium chloride shares its chlorine lone pair of electrons with the aluminium atom for both to be stable. This type of bond exists only in vapour phase after aluminium chloride sublimes.

Diagram method 1

Diagram method 2

A dative/coordinate bond is by convention represented by an arrow (→) heading from the donor of the shared pair of electrons.

Below is the representation of molecules in the above examples;

a)Ammonium ion.

H
H− N→H
H

b)Phosphine ion

H
H− P→H
H

c)Hydroxonium ion

H− O→H
H

d)Carbon(II) oxide

O→C

d) Aluminium(III)chloride

Cl Cl Cl

Al Al

Cl Cl Cl

(ii)IONIC/ELECTROVALENT BOND

An ionic/electrovalent bond is **extreme** of a covalent bond.

During ionic/electrovalent bonding there is complete **transfer** of valence electrons to one electronegative atom from an electropositive atom.

All metals are electropositive and easily/readily donate/lose their valence electrons.

All non-metals are electronegative and easily/readily gain/acquire extra electrons.

Ionic/electrovalent bonding therefore mainly involves transfer of electrons from metal/metallic radical to non-metallic radical.

When an electropositive atom donates /loses the valence electrons, it forms a positively charged **cation** to attain stable octet/duplet.

When an electronegative atom gains /acquires extra valence electrons, it forms a negatively charged **anion** to attain stable octet/duplet**.**

The **electrostatic attraction force** between the **stable** positively charged cation and the stable negatively charged anion with **opposite** charges constitute the ionic bond.

Like in covalent/dative/coordinate bonding, only the outer energy level electrons take part in the formation of ionic/electrovalent bond

Like in covalent/dative/coordinate bonding, the **more** electrons taking part / involved in the formation of ionic/electrovalent bond, the **stronger** the ionic /electrovalent bond.

Illustration of ionic /electrovalent bond

a)Sodium chloride(NaCl)

Sodium chloride(NaCl) is formed when a sodium atom donate its outer valence electrons to chlorine atom for both to attain stable octet:

Symbol of atoms/elements taking part in bonding	Na	Cl
Number of protons/electrons	11	17
Electron configuration/structure	2:8:1	2:8:7
Number of electron in outer energy level	11	7
Number of electrons donated and gained to be stable	1	1
New electron configuration/structure	2:8:	2:8:
Symbol of cation/anion after bonding	Na^+	Cl^-

Diagram

b)Magnesium chloride($MgCl_2$)

Magnesium chloride ($MgCl_2$) is formed when a magnesium atom donate its two outer valence electrons to chlorine atoms. Two chlorine atoms are required to gain each one electron. All the ions (cations and anions) attain stable octet:

Symbol of atoms/elements taking part in bonding	Mg	Cl
Number of protons/electrons	11	17
Electron configuration/structure	2:8:2	2:8:7
Number of electron in outer energy level	2	7
Number of electrons donated and gained to be stable	2	1
New electron configuration/structure	2:8:	2:8:
Symbol of cation/anion after bonding	Mg^{2+}	Cl^{-}

Diagram

c)Lithium oxide(Li_2O)

Lithium oxide(Li_2O)is formed when a Lithium atom donate its outer valence electrons to Oxygen atom. Two Lithium atoms are required to donate/lose each one electron and attain stable duplet. Oxygen atom acquires the two electrons and attain stable octet:

Symbol of atoms/elements taking part in bonding	Li	O
Number of protons/electrons	3	8
Electron configuration/structure	2:1	2:6
Number of electron in outer energy level	1	6
Number of electrons donated and gained to be stable	1	2
New electron configuration/structure	2:	2:8:
Symbol of cation/anion after bonding	Li^{+}	O^{2-}

Diagram

d)Aluminium(III) oxide(Al_2O_3)

Aluminium(III) oxide(Al_2O_3)is formed when a Aluminium atom donate its three outer valence electrons to Oxygen atom. Two Aluminium atoms are required to donate/lose each three electron and attain stable octet. Three Oxygen atoms gain/ acquire the six electrons and attain stable octet:

Symbol of atoms/elements taking part in bonding	Al	O
Number of protons/electrons	13	8
Electron configuration/structure	2:8:3	2:6
Number of electron in outer energy level	3	6
Number of electrons donated and gained to be stable	3	2
New electron configuration/structure	2:8:	2:8:
Symbol of cation/anion after bonding	Al^{3+}	O^{2-}

Diagram

e)Calcium oxide(CaO)

Calcium oxide(CaO)is formed when a Calcium atom donate its two outer valence electrons to Oxygen atom. Both attain stable octet:

Symbol of atoms/elements taking part in bonding	Ca	O
Number of protons/electrons	20	8
Electron configuration/structure	2:8:8:2	2:6
Number of electron in outer energy level	2	6
Number of electrons donated and gained to be stable	2	2
New electron configuration/structure	2:8:8:	2:8:
Symbol of cation/anion after bonding	Ca^{2+}	O^{2-}

Diagram

Some compounds can be formed from ionic/electrovalent, covalent and dative/coordinate bonding within their atoms/molecules:

a)Formation of ammonium chloride:

Ammonium chloride is formed from the reaction of ammonia gas and hydrogen chloride gas. Both ammonia and hydrogen chloride gas are formed from covalent bonding. During the reaction of ammonia and hydrogen chloride gas to form Ammonium chloride;

-ammonia forms a dative/coordinate bond with electron deficient H^+ ion from Hydrogen chloride to form ammonium ion(NH_4^+)ion.

-the chloride ion Cl^- and ammonium ion(NH_4^+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.
Diagram

b) Dissolution/dissolving of hydrogen chloride:
Hydrogen chloride is formed when hydrogen and chlorine atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When hydrogen chloride gas is dissolved in water;

-water molecules forms a dative/coordinate bond with electron deficient H^+ ion from Hydrogen chloride to form hydroxonium ion(H_3O^+)ion.

-the chloride ion Cl^- and hydroxonium ion(H_3O^+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.
Diagram

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c)Dissolution/dissolving of ammonia gas:
Ammmonia gas is formed when hydrogen and Nitrogen atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When Ammonia gas is dissolved in water;

-ammonia forms a dative/coordinate bond with electron deficient H^+ ion from a water molecule to form ammonium ion(NH_4^+)ion.

-the hydroxide ion OH^- and ammonium ion(NH_4^+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

(iii)**METALLIC BOND**

A metallic bond is formed when metallic atoms **delocalize** their outer electrons inorder to be stable.

Metals delocalize their outer electrons to form positively charged cation .

The electrostatic attraction force between the metallic cation and the negatively charged electrons constitute the metallic bond.

The more delocalized electrons the stonger the metallic bond.

Illustration of ionic /electrovalent bond

a) Sodium (Na) is made of one valence electron. The electron is donated to form Na^+ ion. The electron is delocalized /free within many sodium ions.

Symbol of atoms/elements taking part in bonding	Na	Na	Na
Number of protons/electrons	11	11	11
Electron configuration/structure	2:8:1	2:8:1	2:8:1
Number of electron in outer energy level	1	1	1
Number of electrons delocalized/free within	1	1	1
New electron configuration/structure	2:8:	2:8:	2:8:
Symbol of cation after metallic bonding	Na^+	Na^+	Na^+

Diagram

(three)Metallic cations **attract**
(three) free/delocalized electrons

b) Aluminium (Al) is made of three valence electron. The three electrons are donated to form Al^{3+} ion. The electrons are delocalized /free within many aluminium ions.

Symbol of atoms/elements taking part in bonding	Al	Al	Al
Number of protons/electrons	13	13	13
Electron configuration/structure	2:8:3	2:8:3	2:8:3
Number of electron in outer energy level	3	3	3
Number of electrons delocalized/free within	3	3	3
New electron configuration/structure	2:8:	2:8:	2:8:
Symbol of cation after metallic bonding	Al^{3+}	Al^{3+}	Al^{3+}

Diagram

(three)Metallic cations **attract**
(nine) free/delocalized electrons

c)Calcium (Ca) is made of two valence electron.The two electrons are donated to form Ca^{2+} ion.The electrons are delocalized /free within many Calcium ions.

Symbol of atoms/elements taking part in bonding	Ca	Ca	Ca

Number of protons/electrons	20	20	20
Electron configuration/structure	2:8:8:2	2:8:8:2	2:8:8:2
Number of electron in outer energy level	2	2	2
Number of electrons delocalized/free within	2	2	2
New electron configuration/structure	2:8:8:	2:8:8:	2:8:8:
Symbol of cation after metallic bonding	Ca^{2+}	Ca^{2+}	Ca^{2+}

Diagram

(three)Metallic cations **attract**
(six) free/delocalized electrons

d) Magnesium (Mg) is made of two valence electron. The two electrons are donated to form Mg^{2+}ion.The electrons are delocalized /free within many Magnesium ions.

Symbol of atoms/elements taking part in bonding	Mg	Mg
Number of protons/electrons	12	12
Electron configuration/structure	2:8:2	2:8:2
Number of electron in outer energy level	2	2
Number of electrons delocalized/free within	2	2
New electron configuration/structure	2:8:	2:8:
Symbol of cation after metallic bonding	Mg^{2+}	Mg^{2+}

Diagram

(two)Metallic cations **attract**
(four) free/delocalized electrons

e)Lithium (Li) is made of one valence electron.The electron is donated to form Li^{+} ion.The electron is delocalized /free within many Lithium ions.ie;

Symbol of atoms/elements taking part in bonding	Li	Li	Li	Li
Number of protons/electrons	3	3	3	3
Electron configuration/structure	2:1	2:1	2:1	2:1
Number of electron in outer energy level	1	1	1	1
Number of electrons delocalized/free within	1	1	1	1
New electron configuration/structure	2:1:	2:1:	2:1:	2:1:
Symbol of cation after metallic bonding	Li^{+}	Li^{+}	Li^{+}	Li^{+}

Diagram

(four)Metallic cations **attract**
(four) free/delocalized electrons

B.CHEMICAL STRUCTURE

Chemical structure is the pattern/arrangement of atoms **after** they have bonded. There are two main types of chemical structures:

(i)simple molecular structure
(ii) giant structures

(i)Simple molecular structure

Simple molecular structure is the pattern formed after atoms of non-metals have **covalently** bonded to form simple molecules.

Molecules are made of atoms joined together by weak intermolecular forces called **Van-der-waals forces.**The Van-der-waals forces hold the **molecules** together while the covalent bonds hold the **atoms** in the molecule.

Illustration of simple molecular structure

a)Hydrogen molecule(H_2)

Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

Illustration of simple molecular structure

a)Hydrogen molecule(H_2)
Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces

b)Oxygen molecule(O_2)
Oxygen gas is made up of strong covalent bonds/**intra**molecular forces between each Oxygen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

Strong intramolecular forces/covalent bond

O=O:::: O=O:::: O=O:::: O=O
: :   : :   : :   : :   : :   : :   **weak** intermolecular
O=O:::: O=O:::: O=O:::: O=O   forces/van-der-waals forces

c)Iodine molecule(I_2)
Iodine solid crystals are made up of strong covalent bonds/**intra**molecular forces between each iodine atom making the molecule.Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

Strong intramolecular forces/covalent bond

I--- I:::: I --- I:::: I --- I:::: I --- I
: :   : :   : :   : :   : :   : :   : :   **weak** intermolecular
I --- I:::: I --- I:::: I --- I:::: I --- I   forces/van-der-waals forces

d)Carbon(IV) oxide molecule(CO_2)
Carbon(IV) oxide gas molecule is made up of strong covalent bonds/**intra**molecular forces between each Carbon and oxygen atoms making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

Strong intramolecular forces/covalent bond

O=C=O:::: O=C=O:::: O=C=O
: :   : :   : :   **weak** intermolecular
O=C=O:::: O=C=O:::: O=C=O   forces/van-der-waals forces

The following are the main characteristic properties of simple molecular structured compounds:

a)**State**
Most simple molecular substances are gases, liquid or liquids or solid that sublimes or has low boiling/melting points at room temperature (25^{o}C) and pressure (atmospheric pressure).
Examples of simple molecular substances include:

-all gases eg Hydrogen, oxygen, nitrogen, carbon (IV) oxide,

-Petroleum fractions eg Petrol, paraffin, diesel, wax,
-Solid non-metals eg Sulphur, Iodine
-Water

b) **Low melting/boiling points**

Melting is the process of **weakening** the intermolecular/ van-der-waal forces/ of attraction between the molecules that holding the substance/compound.

Note;

(i)Melting and boiling does not involve weakening/breaking the strong intramolecular force/covalent bonds holding the atoms **in** the molecule.

(ii) Melting and boiling points increase with increase in atomic radius/size of the atoms **making** the molecule as the intermolecular forces / van-der-waal forces of attraction between the molecules increase. e.g.

Iodine has a higher melting/boiling point than chlorine because it has a higher /bigger atomic radius/size than chlorine, making the molecule to have stronger intermolecular force/ van-der-waal forces of attraction between the molecules than chlorine. Iodine is hence a solid and chlorine is a gas.

(c)Insoluble in water/soluble in organic solvents

Polar substances dissolve in polar solvents. Water is a polar solvent .Molecular substances do not thus dissolve in water because they are non-polar. They dissolve in non-polar solvents like methylbenzene, benzene, tetrachloromethane or propanone.

d)Poor conductors of heat and electricity

Substances with free mobile **ions** or free mobile/delocalized **electrons** conduct electricity. Molecular substances are poor conductors of heat/electricity because their molecules have no free mobile ions/electrons. This makes them very good **insulators**.

Hydrogen bonds

A hydrogen bond is an intermolecular force of attraction in which a very electronegative atom attracts hydrogen atom of another molecule.

The most electronegative elements are Fluorine, Oxygen and Nitrogen .Molecular compounds made up of these elements usually have hydrogen bonds.

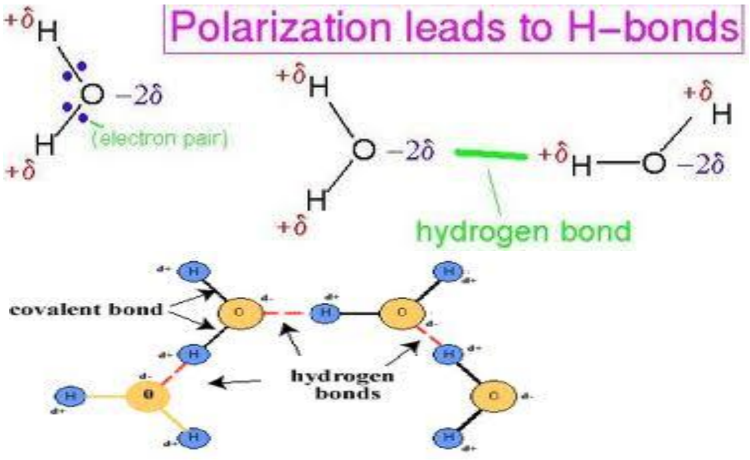
Hydrogen bonds are **stronger** than van-der-waals forces but **weaker** than covalent bonds. Molecular compounds with hydrogen bonds thus have higher melting/boiling points than those with van-der-waals forces.

Illustration of Hydrogen bonding

a)Water molecule

During formation of covalent bond, the oxygen atom attract/pull the shared electrons more to itself than Hydrogen creating partial negative charges(δ^-)in Oxygen and partial positive charges(δ^+)in Hydrogen.

Two molecules **attract** each other at the partial charges through Hydrogen bonding.



The hydrogen bonding in water makes it;

(i)a **liquid** with higher boiling and melting point than simple molecular substances with higher molecular mass. e.g. Hydrogen sulphide as in the table below;

Influence of H-bond in water (H_2O) in comparison to H_2S

Substance	**Water/ H_2O**	**Hydrogen sulphide/ H_2S**
Relative molecular mass	18	34
Melting point(°C)	0	-85
Boiling point(°C)	100	-60

(ii)have higher **volume** in solid (ice) than liquid (water) and thus ice is less dense than water. Ice therefore floats above liquid water.

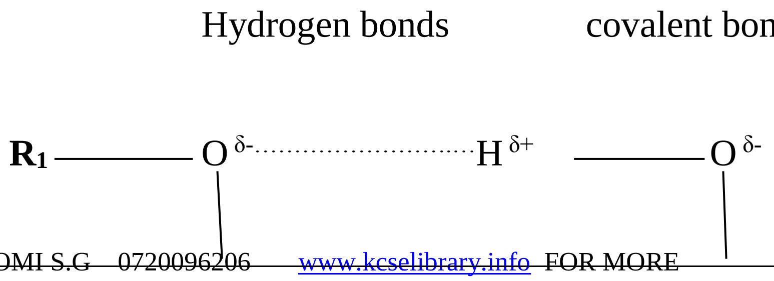
b)Ethanol molecule

Like in water, the oxygen atom attracts/pulls the shared electrons in the covalent bond more to itself than Hydrogen.

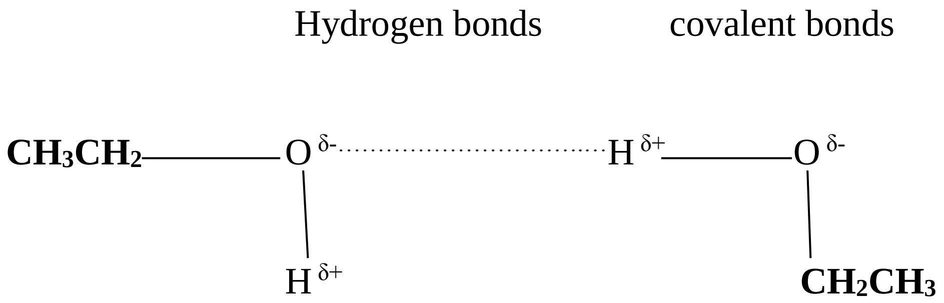
This creates a partial negative charge ($^{\delta-}$) on oxygen and partial positive charge($^{\delta+}$) on hydrogen.

Two ethanol molecules attract each other at the partial charges through Hydrogen bonding forming a **dimmer**.

A dimmer is a molecule formed when two molecules join together as below:



R_1 and R_2 are extensions of the molecule.
For ethanol it is made up of CH_3CH_2 – to make the structure:

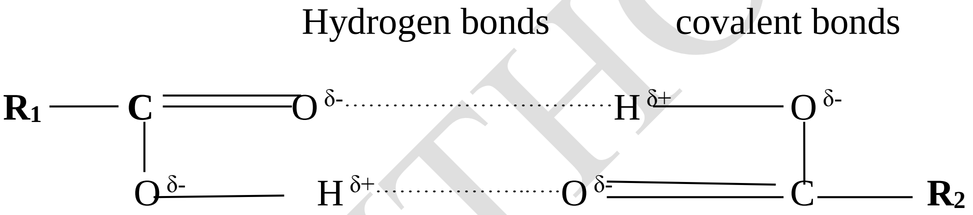


b)Ethanoic acid molecule

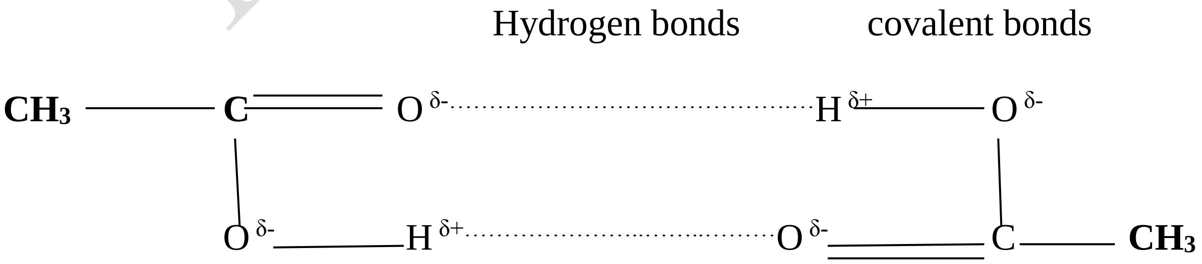
Like in water and ethanol above, the oxygen atom attracts/pulls the shared electrons in the covalent bond in ethanoic acid more to itself than Hydrogen.

This creates a partial negative charge ($^{\delta-}$)on oxygen and partial positive charge($^{\delta+}$) on hydrogen.

Two ethanoic acid molecules attract each other at the partial charges through Hydrogen-bonding forming a **dimer.**



R_1 and $_2$ are extensions of the molecule.
For ethanoic acid the extension is made up of CH_3 – to make the structure;



Ethanoic acid like ethanol exists as a dimer.

Ethanoic acid has a **higher** melting/boiling point than ethanol .This is because ethanoic acid has **two/more** hydrogen bond than ethanol.

d) Proteins and sugars in living things also have multiple/**complex** hydrogen bonds in their structures.

(ii) Giant structure

This is the pattern formed after substances /atoms /ions bond to form a long chain network. Giant structures therefore extend in all directions to form a pattern that continues **repeating** itself.

There are **three** main giant structures.

a) giant covalent/atomic structure
b)giant ionic structure
c)giant metallic structure

a) giant covalent/atomic structure

Giant covalent/atomic structure is the pattern formed after atoms have covalently bonded to form long chain pattern consisting of indefinite number of atoms covalently bonded together. The strong covalent bonds hold all the atoms together to form a very well packed structure.

Examples of substances with giant covalent/atomic structure include:

(i) carbon-**diamond**
(ii) carbon-**graphite**
(iii)silicon
(iv) silicon(IV) oxide/sand

Carbon-graphite and carbon-diamond are **allotropes** of carbon.

Allotropy is the existence of an element in more than one stable physical **form** at the same temperature and pressure.

Allotropes are atoms of the same element existing in more than one stable physical form at the same temperature and pressure.

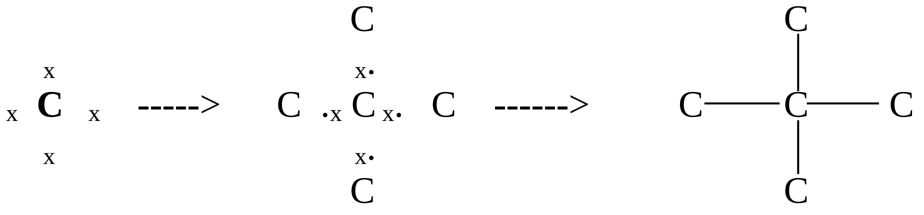
Other elements that exhibit/show allotropy include;

-Sulphur as **monoclinic** sulphur and **rhombic** sulphur
-Phosphorus as **white** phosphorus and **red** phosphorus

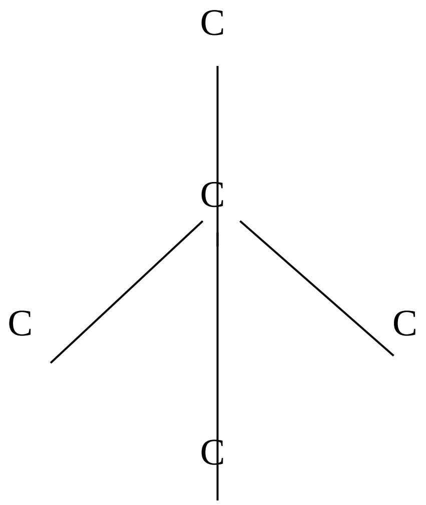
The structure of carbon-diamond

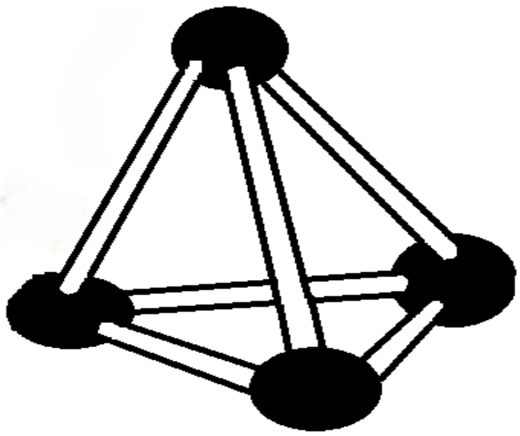
Carbon has four valence electrons. The four valence electrons are used to form covalent bonds.

During the formation of diamond, one carbon atom covalently bond with four other carbon atoms.

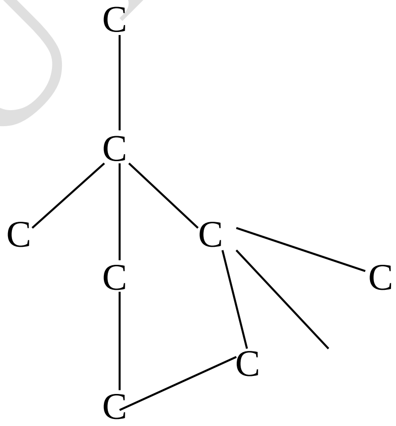


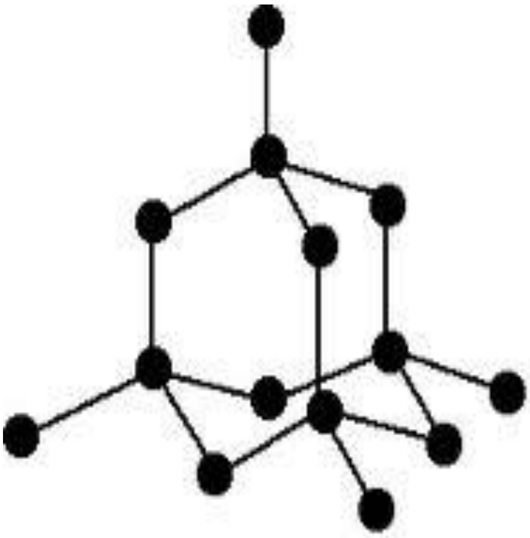
After the bonding, the atoms rearrange to form a regular **tetrahedral** in which **one** carbon is in the centre while **four** are at the apex/corners.





This pattern repeats itself to form a long chain number of atoms covalently bonded together indefinitely. The pattern is therefore called **giant tetrahedral structure.** It extends in all directions where one atom of carbon is always a centre of four others at the apex/corner of a regular tetrahedral.





The giant tetrahedral structure of carbon-diamond is very well/closely packed and joined/bonded together by strong covalent bond.

This makes carbon-diamond to have the following properties:

a) High melting/boiling point.

The giant tetrahedral structure is very well packed and joined together by strong covalent bonds.

This requires a lot of energy/heat to weaken for the element to melt and break for the element to boil.

b) High density.

Carbon diamond is the **hardest** known **natural** substance.

This is because the giant tetrahedral structure is a very well packed pattern/structure and joined together by strong covalent bonds.

This makes Carbon diamond be used to make **drill** for drilling boreholes/oil wells.

The giant tetrahedral structure of carbon diamond is a very closely packed pattern /structure such that heat transfer by **conduction** is possible. This makes carbon diamond a **good thermal** conductor.

c) Poor conductor of electricity.

Carbon-diamond has **no free/delocalized electrons** within its structure and thus do not conduct electricity.

d) Insoluble in water.

Carbon-diamond is insoluble in water because it is non-polar and do not bond with water molecules.

e) Is abrasive/Rough.

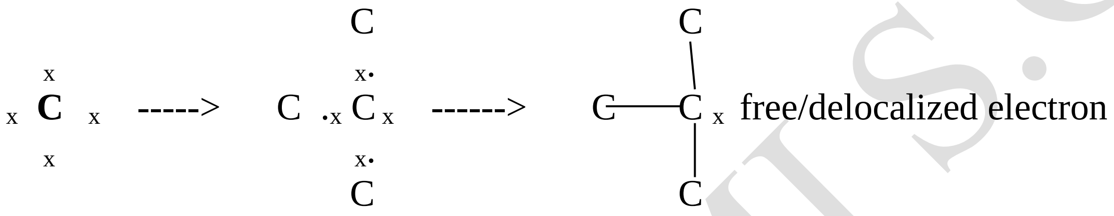
The edges of the closely well packed pattern/structure of Carbon-diamond make its surface rough/abrasive and thus able to smoothen /cut metals and glass.

f) Have characteristic luster.

Carbon-diamond has a high **optical dispersion** and thus able to disperse light to different colours .This makes Carbon-diamond one of the most popular **gemstone** for making **jewellery**.

The structure of carbon-graphite

During the formation of graphite, one carbon atom covalently bond with **three** other carbon atoms leaving **one** free/delocalized electron.



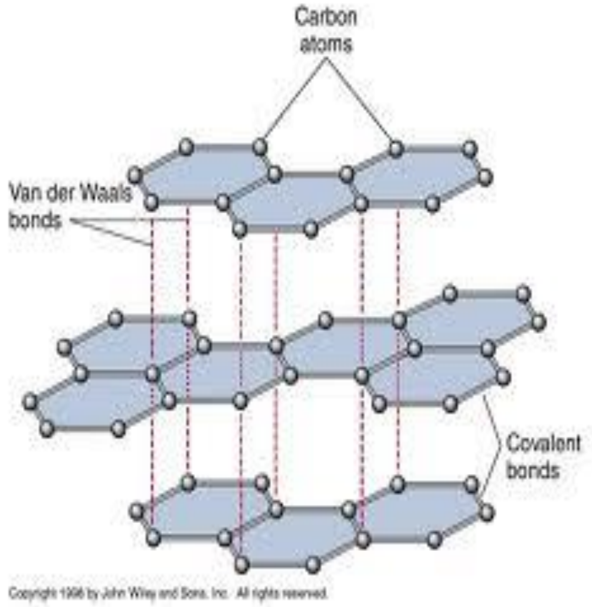
After the bonding, the atoms rearrange and join together to form a regular **hexagon** in which **six** carbon atoms are at the **apex/**corners**.**

The regular hexagon is joined to another in layers on the **same surface** by van-der-waals forces.

Each layer extends to form a plane in all directions.

The fourth valence electron that does not form covalent bonding is free/mobile /delocalized within the layers.

This structure/pattern is called **giant hexagonal planar** structure.



The giant hexagonal planar structure of carbon-graphite is closely packed and joined/bonded together by strong covalent bonds. This makes carbon-graphite to have the following properties:

a) High melting/boiling point.

The giant hexagonal planar structure of carbon-graphite is well packed and joined together by strong covalent bonds.

This requires a lot of energy/heat to **weaken** for the element to melt and **break** for the element to boil.

b) Good conductor of electricity.

Carbon-graphite has **free/delocalized 4th valence electrons** within its structure and thus conducts electricity.

c) Insoluble in water.

Carbon-graphite is insoluble in water because it is **non-polar** and do not bond with water molecules.

d) Soft.

Layers of giant hexagonal planar structure of carbon graphite are held together by van-der-waals forces.

The van-der-waals forces easily break when pressed and reform back on releasing/reducing pressure/force thus making graphite soft.

e) Smooth and slippery.

When pressed at an **angle** the van-der-waals forces easily break and slide over each other making graphite soft and slippery.

It is thus used as a **dry** lubricant instead of oil.

f)Some uses of carbon-graphite.

1. As a dry lubricant- carbon graphite is smooth and slippery and thus better lubricant than oil.Oil **heat up** when reducing friction.
2. Making Lead-pencils- When pressed at an **angle** on paper the van-der-waals forces easily break and slide smoothly over contrasting background producing its characteristic black background.
3. As moderator in nuclear reactors to reduce the rate of decay/disintegration of radioactive nuclides/atoms/isotopes.
4. As electrode in dry/wet cells/battery- carbon graphite is inert and good conductor of electricity. Current is thus able to move from one electrode/terminal to the other in dry and wet cells/batteries. Carbon graphite is also very **cheap**.

b) giant ionic structure

Giant ionic structure is the pattern formed after **ions** have bonded through ionic/electrovalent bonding to form a long chain consisting of **indefinite** number of ions.

The strong ionic/electrovalent bond holds all the **cations** and **anions** together to form a very well packed structure.

Substances with giant ionic structure are mainly crystals of salts e.g. sodium chloride, Magnesium chloride, Sodium iodide, Potassium chloride, copper (II) sulphate(VI).

The structure of sodium chloride

Sodium chloride is made up of sodium (Na^+) and chloride (Cl^-)ions.

Sodium (Na^+) ion is formed when a sodium atom donate /loose/donate an electron. Chloride (Cl^-) ion is formed when a chlorine atom gain /acquire an extra electron from sodium atom.

Many Na^+ and Cl^- ions then rearrange such that **one** Na^+ ion is surrounded by **six** Cl^- ions and one Cl^- ion is surrounded by six Na^+ ions.

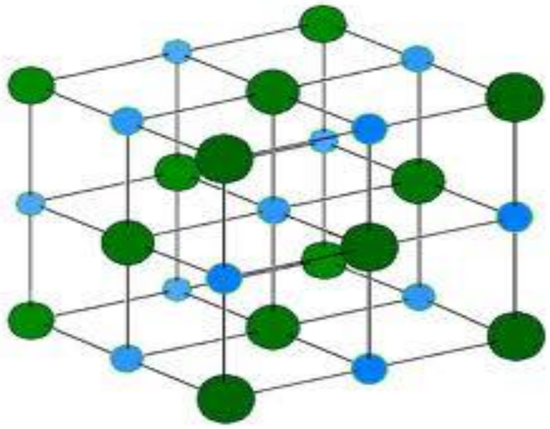
The pattern formed is a **giant cubic structure** where Cl^- ion is sand witched between Na^+ ions and the same to Na^+ ions.

This pattern forms a **crystal**.

A crystal is a solid form of a substance in which particles are arranged in a definite pattern regularly repeated in three dimensions.

The structure of sodium chloride

The giant cubic structure/crystal of sodium chloride is as below;



The giant cubic structure/crystal of sodium chloride is very well packed and joined by strong ionic/electrovalent bonds. This makes sodium chloride and many ionic compounds to have the following properties:

a) Have high melting /boiling points.

The giant cubic lattice structure of sodium chloride is very **closely** packed into a crystal that requires a lot of energy/heat to **weaken** and melt/boil. This applies to all crystalline ionic compounds.

b) Are good conductors of electricity in molten and aqueous state but **poor conductor of electricity in solid.**

Ionic compounds have **fused** ions in solid crystalline state.

On heating and dissolving in water, the crystal is broken into free mobile ions (Na^+ and Cl^- ions).
The free mobile **ions** are responsible for **conduct**ing electricity in ionic compounds in molten and aqueous states.

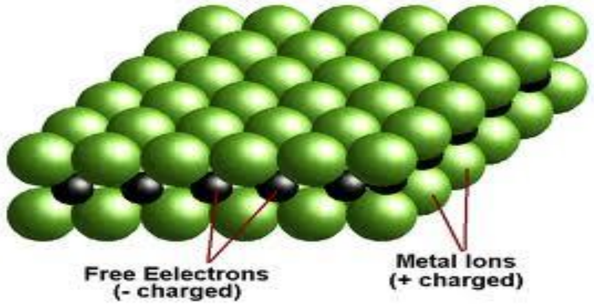
c)Soluble in water
Ionic compounds are **polar** and dissolve in **polar** water molecules.
On dissolving, the crystal breaks to free the fused ions which are then surrounded by water molecules.

b) giant metallic structure

This is the pattern formed after metallic atoms have bonded through metallic bond.
The pattern formed is one where the metallic cations rearrange to form a cubic structure.
The cubic structure is bound together by the free delocalized electrons that move freely within.
The **more** delocalized electrons, the **stronger** the metallic bond.

The structure of sodium and aluminium.

Sodium has one valence electrons.
Aluminium has three valence electrons.
After delocalizing the valence electrons ,the metal cations (Na^+ and Al^{3+}) rearrange to the apex /corners of a regular cube that extend in all directions.
The delocalized electrons remain free and mobile as shown below:



The giant cubic structure makes metals to have the following properties:

a) Have high melting/boiling point
The giant cubic structure is very well packed and joined/bonded together by the free delocalized electrons.
The more delocalized electrons the higher the melting/boiling point.
The larger/bigger the metallic cation ,the weaker the packing of the cations and thus the lower the melting/boiling point. e.g.

(i) Sodium and potassium have both one valence delocalized electron.

Atomic radius of potassium is larger/bigger than that of sodium and hence less well packed in its metallic structure.

Sodium has therefore a higher melting/boiling point than potassium.

(ii) Sodium has one delocalized electron.

Aluminium has three delocalized electrons.

Atomic radius of sodium is larger/bigger than that of aluminium and hence less well packed in its metallic structure.

Aluminium has therefore a higher melting/boiling point than sodium because of the smaller well packed metallic (Al^{3+})ions and bonded/joined by more/three delocalized electrons.

The table below shows the **comparative** melting/boiling points of some metals:

Metal	**Electronic structure**	**Atomic radius(nM)**	**Melting point(°C)**	**Boiling point(°C)**
Sodium	2:8:1	0.155	98	890
Potassium	2:8:8:1	0.203	64	774
Magnesium	2:8:2	0.136	651	1110
Aluminium	2:8:3	0.125	1083	2382

b) Good electrical and thermal conductor/electricity.

All metals are good conductors of heat and electricity including Mercury which is a **liquid**. The mobile delocalized electrons are free within the giant metallic structure to move from one end to the other transmitting heat/electric current.

The more delocalized electrons the better the thermal/electrical conductivity.

High temperatures/heating lowers the thermal/electrical conductivity of metals because the delocalized electrons vibrate and move randomly hindering transfer of heat

From the table above:

Compare the electrical conductivity of;

(i)**Magnesium and sodium**

Magnesium is a better conductor than sodium.

Magnesium has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

(ii)**Potassium and sodium**

Potassium is a better conductor than sodium.

Potassium has bigger/larger atomic radius than sodium. The delocalized electrons are less attracted to the nucleus of the atom and thus more free /mobile and thus better the electrical conductor.

c) Insoluble in water

All metals are insoluble in water because they are non polar and thus do not bond with water. Metals higher in the reactivity/electrochemical series like; Potassium, sodium, Lithium and Calcium **reacts** with **cold** water producing hydrogen gas and forming an **alkaline** solution of their **hydroxides**.ie

$$2K(s) \quad + \quad 2H_2O(l) \quad -> \quad 2KOH(aq) \quad + \quad H_2(g)$$

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

Heavy metal like Magnesium, Aluminium, Iron, Zinc and Lead **react** with **steam/water vapour** to produce hydrogen gas and form the corresponding **oxide**.

$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$
$Fe(s) + H_2O(g) \rightarrow FeO(s) + H_2(g)$
$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$
$Pb(s) + H_2O(g) \rightarrow PbO(s) + H_2(g)$
$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$

Metals **lower** in the reactivity/electrochemical series than hydrogen like; copper, Mercury, Gold Silver and Platinum **do not** react with water/vapour.

d) Shiny metallic-lustre

All metals have a shiny **grey** metallic luster <u>except</u> copper which is **brown.**
When exposed to sunlight, the delocalized electrons gain energy, they vibrate on the metal surface scattering light to appear shiny.
With time, most metals corrode and are covered by a layer of the metal oxide.
The delocalized electrons are unable to gain and scatter light and the metal surface tarnishes/become dull.

e) Ductile and malleable

All metals are malleable (can be made into thin **sheet**) and ductile (can be made into **wire**.
When beaten/hit/pressed **lengthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a sheet.
When beaten/hit/pressed **lengthwise and bredthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a wire/thin strip.

f) Have high tensile strength

Metals are not brittle. The free delocalized electrons bind the metal together when it is bent /coiled at any angle.
The meta thus withstand stress/coiling

g) Form alloys

An alloy is a uniform mixture of two or more metals.
Some metals have spaces between their metallic cations which can be occupied by another metal cation with smaller atomic radius.
Common alloys include:

Brass(Zinc and Copper alloy)
Bronze(Copper and Tin alloy)
German silver

Summary of Bonding and structure

	Simple molecular structure	**Giant covalent /atomic structure**	**Giant ionic structure**	**Giant metallic structure**
(i)Examples	I_2,S_8,HCl,O_2,CH_4	Graphite,diamond Si,SiO_2	NaCl, KCl, CaO,$CuSO_4$	Na,Fe,Cr,Hg,K
Constituent particles making	molecules	Atoms (of non-metals)	Ions (cation and anions)	Atoms (of metals)

structure				
Type of substance	Non-metal element/non-metal molecule/non-metal compound(electronegative elements)	Group IV non-metals and some of their oxides	Metal-non metal compounds(compounds of electropositive and electronegative compounds)	Metallic elements (with low electonegativity and high electropositivity)
Bonding in solid state	-Strong covalent bonds hold atoms together within separate molecules (intramolecular forces) -Weak van-der-waals forces hold separate molecules together (intermolecular forces)	Atoms are linked through the whole structure by very strong covalent bonds.	Electrostatic attraction of cations and anions link the whole structure through strong ionic bond.	Electrostatic attraction of outer mobile electrons for positive nuclei binds atoms together though metallic bond
Properties (i) Volatility	-Highly volatile with low melting/boiling point -Low latent heat of fusion/vaporization	-Non volatile with very high melting/boiling points -Low latent heat of fusion / vaporization	-Non volatile with very high melting/boiling points -Low latent heat of fusion / vaporization	-Non volatile with very high melting/boiling points -Low latent heat of fusion / vaporization
(ii) State at room temperature /pressure	Usually gases,volatile liquids or solids that sublimes	solids	solids	Solids except Mercury(liquid)
(iii) Hardness	Soft and brittle(low tensile strength)	Hard and brittle(low tensile strength)	Hard and brittle(low tensile strength)	Hard, malleable, ductile and have high tensile strength

(iv) Thermal /electrical conductivity	Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but some dissolve and react to form electrolytes e.g. Hydrogen chloride and ammonia gases.	Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but -Carbon-graphite is a good electrical conductor while -Carbon-diamond is a good thermal conductor.	Poor thermal and electrical conductor when solid. Good thermal and electrical conductor in liquid/molten and aqueous states when the ions are not fused	Good thermal and electrical conductor in solid and liquid/molten states due to the free mobile /delocalized electrons
(v) Solubility	Insoluble in polar solvents e.g. Water Soluble in non-polar solvents e.g. tetrachloromethane, benzene, methylbenzene	Insoluble in all solvents	Soluble in polar solvents e.g. Water Insoluble in non-polar solvents e.g. tetrachloromethane, benzene, methylbenzene	Insoluble in polar/non-polar colvents. -Some react with polar solvents -Some metal dissolve in other metals to form alloys e.g. Brass is formed when Zinc dissolve in copper.

C. PERIODICITY OF BONDING AND STRUCTURE

The periodic table does not classify elements as metals and non-metals. The table arranges them in terms of atomic numbers.

However, based on structure and bonding of the elements in the periodic table;

(i)-the top right hand corner of about twenty elements are non-metals

(ii)-left of each non-metal is an element which shows characteristics of both metal and non-metal.

These elements are called **semi-metals/metalloids**. They include Boron, silicon, Germanium, Arsenic, and Terullium

(iii)-all other elements in the periodic table are metal.

(iv)-Hydrogen is a non-metal with metallic characteristic/property of donating/losing outer electron to form cation/H^+ ion.

(v) –bromine is the only known natural liquid non-metal element at room temperature and pressure.

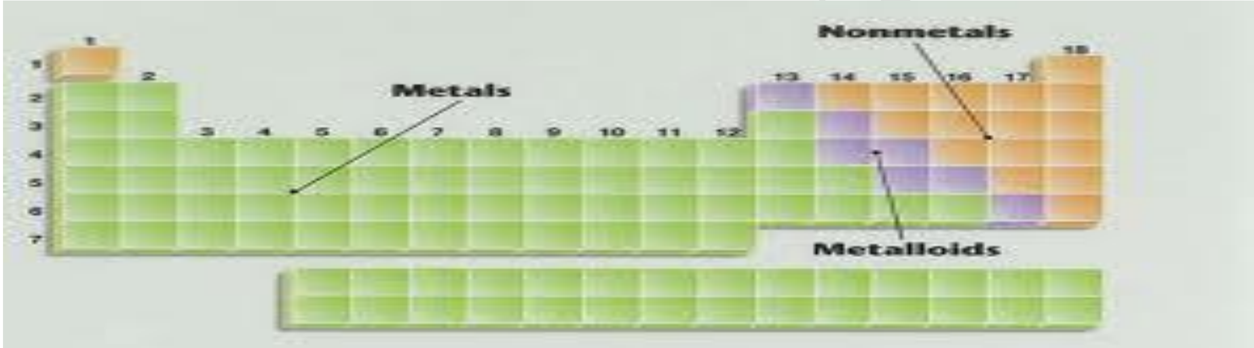
(vi) –mercury is only known natural liquid metal element at room temperature and pressure.

(vii) Carbon-graphite is a semi metals/metalloids. Carbon-diamond is a pure non-metal yet both are allotropes of carbon (same element)

a) Sketch of the periodic table showing metals ,metalloid and non-metals

Metals **Metalloids** *Non-metals*

H								*He*
Li	Be		**B**	**C**	*N*	*O*	*F*	*Ne*
Na	Mg		Al	**Si**	*P*	*S*	*Cl*	*Ar*
K	Ca	Transition metals	Ga	Ge	**As**	*Se*	*Br*	*Kr*
Rb	Sr		In	Sn	Sb	**Te**	*I*	*Xe*
Cs	Ba		Tl	Pb	Bi	Po	**At**	*Rn*
Fr	Ra							



b)Periodicity in the physical properties of elements across period 2 and 3

Study table I and II below:

Table I(period 2)

Property	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Melting point(°C)	180	1280	2030	3700 (graphite) 3550 (diamond)	-210	-219	-220	-250
Boiling point(°C)	1330	2480	3930	Graphite sublimes 4830 (diamond)	-200	-180	-190	-245
Density at room	0.50	1.85	2.55	2.25 (graphite)	0.81	0.14	0.11	0.021

temperature (gcm^{-3})				3.53 (diamond)				
Type of element	Metal	Metal	Metal	Metalloid	Non-metal	Non-metal	Non-metal	Non-metal
Chemical structure	Giant metallic	Giant metallic	Giant atomic/ covalent	Giant atomic/ covalent	Simple molecula or molecule/ N_2	Simple molecula or molecules /O_2	Simple molecula or molecule/F_2	Simple molecula or molecule/Ne
State at room temperature	Solid	Solid	Solid	Solid	gas	gas	gas	gas
Electron structure	2:1	2:2	2:3	2:4	2:5	2:6	2:7	2:8
Valency	1	2	3	4	3	2	1	-
Formular of ion	Li^+	Be^{2+}	B^{3+}	-	N^{3-}	O^{2-}	F^-	-

Table II (period 3)

Property	**Na**	**Mg**	**Al**	**Si**	**P(white)**	**S(Rhombic)**	**Cl**	**Ar**
Melting point(°C)	98	650	660	1410	44	114	-101	-189
Boiling point(°C)	890	1120	2450	2680	280	445	-34	-186
Density at room temperature (gcm^{-3})	0.97	1.74	2.70	2.33 (graphite) 3.53 (diamond)	1.82	2.07	0.157	0.011
Type of	Metal	Metal	Metal	Metalloi	Non-	Non-metal	Non-	Non-

element				d	metal		metal	metal
Chemical structure	Giant metallic	Giant metallic	Giant metallic	Giant atomic/ covalent	Simple molecula or molecule/ P_4	Simple molecula or molecules /S_8	Simple molecula or molecule/Cl_2	Simple molecula or molecule /Ar
State at room temperature	Solid	Solid	Solid	Solid	Solid	Solid	gas	gas
Electron structure	2:8:1	2:8:2	2:8:3	2:8:4	2:8:5	2:8:6	2:8:7	2:8:8
Valency	1	2	3	4	3	2	1	-
Formular of ion	Na^+	Mg^{2+}	Al^{3+}	-	P^{3-}	S^{2-}	Cl^-	-

From table I and II above:

1. Explain the trend in atomic radius along /across a period in the periodic table

Observation

Atomic radius of elements in the same period decrease successively across/along a period from left to right.

Explanation

Across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level.

Increase in nuclear charge increases the effective nuclear attraction on the outer energy level pulling it closer to the nucleus successively across the period .e.g.

(i)From the table 1and 2 above, atomic radius of Sodium (0.157nM) is higher than that of Magnesium(0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus than sodium.

(ii)The rate of decrease in the atomic radius become smaller as the atom become heavier e.g.

Atomic radius of Magnesium from sodium falls by(0.157nM- 0.137nM) =**0.02**

Atomic radius of Chlorine from sulphur falls by(0.104nM- 0.099nM) =**0.005**

This is because gaining/adding one more proton to 11 already present cause greater proportional change in nuclear attraction power to magnesium than gaining/adding one more proton to 16 already present in sulphur to chlorine.

(iii)Period 3 elements have more energy levels than Period 2 elements. They have therefore bigger/larger atomic radius/size than corresponding period 2 elements in the same group.

2.Explain the trend in ionic radius along/across a period in the periodic table

Observation

Ionic radius of elements in the same period decrease successively across/along a period from left to right for the first three elements then increase drastically then slowly successively decrease.

Explanation

Across/along the period from left to right elements change form electron donors/losers (**reducing** agents) to electron acceptors (**oxidizing** agents).

(i)An atom form stable ion by either gaining/acquiring/ accepting extra electron or donating/losing outer electrons.

(ii)Metals form stable ions by donating/losing **all** the outer energy level electrons and thus **also** the outer energy level .i.e.

-Sodium ion has one less energy level than sodium atom. The ion is formed by sodium atom donating/losing (all) the outer energy level electron and thus also the outer energy level making the ion to have smaller ionic radius than atom.

(iii)Ionic radius therefore decrease across/along the period from Lithium to Boron in period 2 and from Sodium to Aluminium in period 3.This is because the number of electrons donated/lost causes increased effective nuclear attraction on remaining electrons /energy levels.

(iv)Non-metals form stable ion by gaining/acquiring/accepting extra electron in the outer energy level. The extra electron/s increases the repulsion among electrons and reduces the effective nuclear attraction on outer energy level. The outer energy level therefore expand/enlarge/increase in order to accommodate the extra repelled electrons .The more electrons gained/accepted/acquired the more repulsion and the more expansion to accommodate them and hence bigger/larger atomic radius. e.g.

-Nitrogen ion has three electrons more than Nitrogen atom. The outer energy level expand/enlarge/increase to accommodate the extra repelled electrons. Nitrogen atom thus has smaller atomic radius than the ionic radius of nitrogen ion.

(v) Ionic radius decrease from group IV onwards from left to right. This because the number of electrons gained to form ion decrease across/along the period from left to right. e.g. Nitrogen ion has bigger/larger ionic radius than Oxygen.

3.Explain the trend in melting and boiling point of elements in a period in the periodic table.

Observation

The melting and boiling point of elements rise up to the elements in Group IV(Carbon/Silicon) along/across the period then continuously falls.

Explanation

Melting/boiling points depend on the packing of the structure making the element and the strength of the bond holding the atoms/molecules together.

Across/along the period (2 and 3) the structure changes from giant metallic, giant atomic/covalent to simple molecular.

(i)For metals, the number of delocalized electrons increases across/along the period and hence stronger metallic bond/structure thus requiring a lot of heat/energy to weaken.

The strength of a metallic bond also depends on the atomic radius/size. The melting /boiling point decrease as the atomic radius/size of metals increase due to decreased packing of larger atoms. e.g.

-The melting /boiling point of Lithium is lower than that of Beryllium because Beryllium has two/more delocalized electrons and hence stronger metallic structure/bond.

- The melting /boiling point of Lithium is higher than that of Sodium because the atomic radius/size Lithium is smaller and hence better packed and hence forms stronger metallic structure/bond.

(ii)Carbon-graphite/carbon-diamond in period 2 and Silicon in period 3 form very well packed giant atomic/covalent structures held together by strong covalent bonds. These elements have therefore very high melting/boiling points.

Both Carbon-graphite/ carbon-diamond have smaller atomic radius/size than Silicon in period 3 and thus higher melting/boiling points due to better/closer packing of smaller atoms in their well packed giant atomic/covalent structures.

(ii)Non-metals from group V along/across the period form simple molecules joined by weak intermolecular /van-der-waals force. The weak intermolecular /van-der-waals force require little energy/heat to weaken leading to low melting/boiling points. The strength of the intermolecular /van-der-waals forces decrease with decrease in atomic radius/ size lowering the melting/boiling points along/across the period (and raising the melting/boiling points down the group).e.g.

-The melting /boiling point of Nitrogen is higher than that of Oxygen. This is because the atomic radius/ size of Nitrogen is higher than that of Oxygen and hence stronger intermolecular /van-der-waals forces between Nitrogen molecules.

-The melting /boiling point of Chlorine is higher than that of Fluorine. This is because the atomic radius/ size of Chlorine is higher than that of Fluorine and hence stronger intermolecular /van-der-waals forces between Chlorine molecules.

(iii)Rhombic sulphur exists as a puckered ring of S_8atoms which are well packed. Before melting the ring break and join to very long chains that entangle each other causing the **unusually** high melting/boiling point of Rhombic sulphur.

(iv)Both sulphur and phosphorus exists as allotropes.
Sulphur exists as **Rhombic**-sulphur and **monoclinic**-sulphur. Rhombic-sulphur is the stable form of sulphur at room temperature and pressure.
Phosphorus exists as **white**-phosphorus and **red**-phosphorus.
White-phosphorus is the stable form of Phosphorus at room temperature and pressure.

4. State and explain the trend in density of elements in a period in the periodic table.
Observation: Density increase upto the elements in group IV then falls across/along the period successively
Explanation:
Density is the mass per unit volume occupied by matter/particles/atoms/molecules of element.

(i)For metals ,the stronger metallic bond and the more delocalized electrons ensure a very well packed giant metallic structure that occupy less volume and thus higher density.
The more the number of delocalized electrons along/across the period, the higher the density. e.g.
(i)Aluminium has a higher density than sodium. This is because aluminium has more /three delocalized electrons than /one sodium thus forms a very well packed giant metallic structure that occupy less volume per given mass/density.

(ii)Carbon-graphite ,carbon-diamond and silicon in group IV form a well packed giant atomic/covalent structure that is continuously joined by strong covalent bonds hence occupy less volume per given mass/density.
Carbon-graphite form a less well packed giant hexagonal planar structure joined by Van-der-waals forces. Its density (2.25gcm^{-3}) is therefore less than that of Carbon-diamond(3.53gcm^{-3}) and silicon(2.33gcm^{-3}).Both diamond and silicon have giant tetrahedral structure that is better packed. Carbon-diamond has smaller atomic radius/size than silicon. Its density is thus higher because of better packing and subsequently higher density. Carbon-diamond is the hardest known natural substance by having the highest density.

(iii)For non-metals, the strength of the intermolecular /van-der-waals forces decreases with decrease in atomic radius/size along/across the period. This decreases the mass occupied by given volume of atoms in a molecule from group VI onwards. e.g.
Phosphorus has a higher atomic radius/size than chlorine and Argon and thus stronger intermolecular/van-der-waals forces that ensure a given mass of phosphorus occupy less volume than chlorine and neon.

5.State and explain the trend in thermal/electrical conductivity of elements in a period in the periodic table.
Observation:
Increase along/across the period from group I, II, and III then decrease in Group IV to drastically decrease in group V to VIII (O).
Explanation

(i)Metals have free delocalized electrons that are responsible for thermal/electrical conductivity.Thermal/electrical conductivity increase with increase in number of delocalized electrons. The thermal conductivity decrease with increase in temperature/heating.
e.g.
Aluminium with three delocalized electrons from each atom in its metallic structure has the highest electrical /thermal conductivity in period 3.

(ii)Carbon-graphite has also free 4th valency electrons that are delocalized within its layers of giant hexagonal planar structure. They are responsible for the electrical conductivity of graphite.

(iii)Silicon and carbon diamond do not conduct electricity but conducts heat. With each atom too close to each other in their very well packed giant tetrahedral structure, heat transfer /radiate between the atoms. The thermal conductivity increase with increase in temperature/heating.

(iv)All other non-metals are poor /non-conductor of heat and electricity. They are made of molecules with no free /mobile delocalized electrons in their structure.

Periodicity of the oxides of elements along/across period 3

The table below summarizes some properties of the oxides of elements in period 3 of the periodic table.

Formular of oxide/ Property	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5 P_4O_6	SO_2 SO_3	Cl_2O_7 Cl_2O
Melting point(°C)	1193	3075	2045	1728	563	-76	-60
Boiling point(°C)	1278	3601	2980	2231	301	-10	-9
Bond type	Ionic	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
Chemical structure	Giant ionic structure	Giant ionic structure	Giant ionic structure	Giant atomic/ covalent	Simple molecula or molecule	Simple molecula or molecules	Simple molecula or molecule
State at room temperature	Solid	Solid	Solid	Solid	Solid	gas	Gas (**Cl_2O_7** is a liquid)
Nature of Oxide	Basic/ alkaline	Basic/ alkaline	Amphotellic oxide	2:8:4	2:8:5	2:8:6	2:8:7
Reaction with water	React to form NaOH /alkaline solution	React to form $MgOH)_2$ /weakly alkaline solution	Don't react with water.	Don't react with water.	React to form H_2PO_4 /weakly acidic solution	-SO_2 react to form H_2SO_3 . H_2SO_3 is quickly oxidized to H_2SO_4 -SO_2 react to form H_2SO_4/ strongly acidic	**-Cl_2O_7 reacts to form $HClO_4$** /weakly acidic solution
Reaction	Reacts	Reacts	Reacts to	No	No	No	No

with dilute acids	to form salt and water	to form salt and water	form salt and water	reaction	reaction	reaction	reaction

1. All the oxides of elements in period 3 except those of sulphur and chlorine are **solids** at room temperature and pressure.
2. Across/along the period, bonding of the oxides changes from **ionic** in sodium oxide magnesium oxide and aluminium oxide (show both ionic and covalent properties) to **covalent** in the rest of the oxides.

3. Across/along the period, the structure of the oxides changes from giant **ionic** structure in sodium oxide, magnesium oxide and aluminium oxide to giant **atomic**/covalent structure in silicon (IV) oxide. The rest of the oxides form simple **molecules**/**molecular** structure.

4. Sodium oxide and magnesium oxide are **basic /alkaline** in nature. Aluminium oxide is **amphotellic** in nature (shows both acidic and basic characteristics). The rest of the oxides are **acidic** in nature.

5. Ionic compounds/oxides have very **high** melting/boiling points because of the strong **electrostatic attraction** joining the giant ionic crystal lattice.
The melting/boiling points increase from sodium oxide to aluminium oxide as the number of electrons involved in bonding increase, increasing the strength of the ionic bond/structure.

6. Silicon (IV) oxide is made of a well packed giant atomic/covalent structure joined by strong covalent bonds.
This results in a solid with very **high** melting/boiling point.

7.Phosphorus (V) oxide, sulphur(IV) oxide/ sulphur (VI) oxide and dichloride heptoxide exist as simple molecules/molecular structure joined by weak van-der-waals/intermolecular forces.
This results in them existing as **low** melting /boiling point solids/gases.

8. Ionic oxide conducts electricity in molten and aqueous states but not in solid.
In solid state the ions are fused/fixed but on heating to molten state and when dissolved in water, the ions are free / mobile.
Sodium oxide, magnesium oxide and aluminium oxide are therefore good conductors in molten and aqueous states.

9. Covalent bonded oxides do not conduct electricity in solid, molten or in aqueous states.
This is because they do not have free / mobile ion. Phosphorus (V) oxide, sulphur(IV) oxide/ sulphur (VI) oxide and dichloride heptoxide are thus non-conductors/insulators.

10. Silicon (IV) oxide is a poor/weak conductor of heat in solid state. This is because it has very closely packed structure for heat to radiate conduct along its structure.

11. Electopositivity decrease across the period while electronegativity increase across the period. The oxides thus become less ionic and more covalent along/across the period.
12.The steady change from giant ionic structure to giant atomic/ covalent structure then simple molecular structure lead to profound differences in the reaction of the oxides with water,acids and alkalis/bases:

(i) Reaction with water

a) Ionic oxides react with water to form alkaline solutions e.g.;

I.Sodium oxide reacts/dissolves in water forming an alkaline solution of sodium hydroxide.

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

II. Magnesium oxide slightly/ slowly reacts/dissolves in water forming an alkaline solution of magnesium hydroxide

Chemical equation: $MgO(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq)$

III. Aluminium oxide does reacts/dissolves in water.

b) Non-metallic oxides are acidic. They react with water to form weakly acidic solutions:

I. Phosphorus (V) oxide readily reacts/dissolves in water forming a weak acidic solution of phosphoric (V) acid.

Chemical equation: $P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$

Chemical equation: $P_2O_5(s) + 3H_2O(l) \rightarrow 2H_3PO_4(aq)$

II. Sulphur (IV) oxide readily reacts/dissolves in water forming a weak acidic solution of sulphuric (IV) acid.

Chemical equation: $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

Sulphur (VI) oxide quickly fumes in water to form concentrated sulphuric (VI) acid which is a strong acid.

Chemical equation: $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

III. Dichlorine oxide reacts with water to form weak acidic solution of chloric(I) acid/hypochlorous acid.

Chemical equation: $Cl_2O(g) + H_2O(l) \rightarrow 2HClO(aq)$

IV. Dichlorine heptoxide reacts with water to form weak acidic solution of chloric(VII) acid.

Chemical equation: $Cl_2O_7(l) + H_2O(l) \rightarrow 2HClO_4(aq)$

c) Silicon (IV) oxide **does not** react with water.

It reacts with hot concentrated alkalis forming silicate (IV) salts. e.g.

Silicon (IV) oxide react with hot concentrated sodium hydroxide to form sodium silicate (IV) salt.

Chemical equation: $SiO_2(s) + 2NaOH(aq) \rightarrow Na_2SiO_3(aq) + H_2O(l)$

(ii) Reaction with dilute acids

a) Ionic oxides react with dilute acids to form salt and water only. This is a **neutralization** reaction. e.g.

Chemical equation: $Na_2O(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$
Chemical equation: $MgO(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$
Chemical equation: $Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$

Aluminium oxide is amphotellic and reacts with **hot concentrated strong** alkalis sodium/potassium hydroxides to form **complex** sodium aluminate(III) and potassium aluminate(III) salt.

Chemical equation: $Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$
Chemical equation: $Al_2O_3(s) + 2KOH(aq) + 3H_2O(l) \rightarrow 2KAl(OH)_4(aq)$

b) Acidic oxides **do not** react with dilute acids.

c)Periodicity of the Chlorides of elements along/across period 3

The table below summarizes some properties of the chlorides of elements in period 3 of the periodic table.

Formular of chloride/ Property	**NaCl**	**$MgCl_2$**	**$AlCl_3$**	**$SiCl_4$**	**PCl_5** **PCl_3**	**SCl_2** S_2Cl_2	**Cl_2**
Melting point(°C)	801	714	Sublimes at 180 **°C**	-70	PCl_5 Sublimes at -94 **°C**	-78	-101

Boiling point(^oC)	1465	1418	423(as **Al_2Cl_6** vapour	57	74(as **P_2Cl_6** Vapour 164 (as **PCl_5)**	decomposes at 59 **^oC**	-34
Bond type	Ionic	Ionic	Ionic/ Covalent/ dative	Covalent	Covalent	Covalent	Covalent
Chemical structure	Giant ionic structure	Giant ionic structure	Molecular / dimerizes	Simple molecula or molecule	Simple molecula or molecule	Simple molecula or molecules	Simple molecula or molecule
State at room temperature	Solid	Solid	Solid	liquid	Liquid **PCl_5 is solid**	liquid	Gas
Nature of Chloride	Neutral	Neutral	Strongly acidic	Strongly acidic	Strongly acidic	Strongly acidic	Strongly acidic
pH of solution	7.0	7.0	3.0	3.0	3.0	3.0	3.0
Reaction with water	Dissolve	Dissolve	-Hydrolysed by water -Acidic hydrogen chloride fumes produced	-Hydrolysed by water -Acidic hydrogen chloride fumes produced	Hydrolysed by water -Acidic hydrogen chloride fumes produced	Hydrolysed by water -Acidic hydrogen chloride fumes produced	Forms HCl and HClO
Electrical conductivity in molten/aqueous state	good	good	poor	nil	nil	nil	nil

1. Sodium Chloride, Magnesium chloride and aluminium chloride are **solids** at room temperature and pressure.
Silicon(IV) chloride, phosphorus(III)chloride and disulphur dichloride are **liquids.** Phosphorus(V)chloride is a **solid**. Both chlorine and sulphur chloride are **gases.**

2. Across/along the period bonding changes from **ionic** in Sodium Chloride and Magnesium chloride to **covalent** in the rest of the chlorides.

3. Anhydrous aluminium chloride is also a molecular compound .Each aluminium atom is covalently bonded to three chlorine atoms.
In vapour/gaseous phase/state two molecules dimerizes to Al_2O_6 molecule through coordinate/dative bonding.

4. Across/along the period the structure changes from giant **ionic** in Sodium Chloride and Magnesium chloride to **simple molecules/molecular structure** in the rest of the chlorides.

5. Ionic chlorides have very high melting /boiling points because of the strong ionic bond/electrostatic attraction between the ions in their crystal lattice.The rest of the chlorides have low melting /boiling points because of the weak van-der-waal /intermolecular forces.

6. Sodium Chloride and Magnesium chloride in molten and aqueous state have free/mobile ions and thus good electrical conductors. Aluminium chloride is a poor conductor. The rest of the chlorides do not conduct because they have no free/mobile **ions**.

7. Ionic chloride form **neutral** solutions with pH =7. These chlorides **ionize**/**dissociate** completely into free cations and anions.i.e;
Sodium Chloride and Magnesium chloride have pH=7 because they are fully/completely ionized/dissociated into free ions.

Chemical equation $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$
Chemical equation $MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$

8 Across/along the period from aluminium chloride, **hydrolysis** of the chloride takes place when reacting/dissolved in water.
Hydrolysis is the reaction of a compound when dissolved in water.

a)Aluminium chloride is hydrolyzed by water to form aluminium **hydroxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(s) + 3HCl(g)$

b)Silicon(IV) chloride is hydrolyzed by water to form silicon(IV)**oxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(g)$

This reaction is highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

c) Both phosphoric (V) chloride and phosphoric (III) chloride are hydrolyzed by water to form phosphoric (V) **acid** and phosphoric (III) **acid** respectively. Fumes of hydrogen chloride gas are produced. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $PCl_5(s) + 4H_2O(l) \rightarrow H_3PO_4(aq) + 5HCl(g)$

Chemical equation PCl_3 (s) + $3H_2O$(l)-> H_3PO_4(aq) + 3HCl(g)

This reaction is also highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

d) Disulphur dichloride similarly hydrolyzes in water to form **yellow** deposits of sulphur and produce a mixture of **sulphur (IV) oxide** and **hydrogen chloride** gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $2S_2Cl_2$ (l) + $2H_2O$(l)-> 3S(s) + SO_2(g) + 4HCl(g)

D. COMPREHENSIVE REVISION QUESTIONS

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

								A
B				G		H	E	C
	J		I	L				
D	***N***						M	

(a) (I) Indicate on the grid the position of an element represented by letter N whose electronic configuration of a divalent cation is 2:8:8 . (1 mark)

(II) Name the bond formed between D and H react. Explain your answer.(2 marks)

Ionic/electrovalent

D is electropositive thus donates two electrons to electronegative H

(III) Write an equation for the reaction between B and water. (1 mark)

Chemical equation 2B (s) + $2H_2O$(l) -> 2BOH(aq) + H_2 (g)

(IV) How do the atomic radii of I and L compare. Explain. (2 marks)

(V) In terms of structure and bonding explain why the oxide of G has lower melting point than oxide of L. (2 marks)

(b) Study the information given below and answer the question that follow.

Formula of compound N	NaCl	$MgCl_2$	Al_2Cl_6	$SiCl_4$	PCl_3	SCl_2
B.P(^{0}C)	1470	1420	Sublimes	60	75	60
M.P(^{0}C)	800	710	At 800^{0}C	-70	90	-80

(I)Why is the formula of aluminium chloride given as Al_2Cl_6 and not $AlCl_3$? (1 mark)

(II) Give two chlorides that are liquid at room temperature. Give a reason for the answer. (2 marks)

(III) Give a reason why Al_2Cl_6 has a lower melting point than $MgCl_2$ although both Al and Mg are metals. (1 mark)

(IV) Which of the chlorides would remain in liquid state for the highest temperature range explain why ? (2 mark)

(Kakamega)

2. 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. (2mks)

(ii) Explain why the melting point of the oxide of R is higher than that of the oxide of S. (2mks)

(iii) Give **two** elements that would react vigorously with each other. Explain your answer. (2mks)

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} ionization energy	2^{nd} ionization energy
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? (1mk)

(ii) What is meant by the term ionization energy? (1mk)

iii) The 2^{nd} ionization energy is higher that the 1^{st} ionization energy of each. Explain (1mk)

(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 burns explosively is produced. Use a simple diagram to illustrate how this gas can be collected during this experiment. (3mks)

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

								A
B			X	G		Z	E	V
	J		I	L		T		
D	***N***						M	

a) Select the most reactive non-metal. (1mk)

b) Write the formula of the compound consisting of

I.D and Z only. (2mk)

II. X and Z

c) Select an element that can form an ion of change +2 (1mk)

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

e) Suggest with reasons a likely pH value of an aqueous solution of the chlorine of:(3mks)

T.

B

X

f) To which chemical family do the following elements belong? (2mk)

J

V

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. (3mks)

4.The grid below shows part of the periodic table study it and answer the questions that follow. The letters do not represent the true symbols.

						A		
	B		**C**		**D**		**E**	
F	**G**							
							H	

\(a) Which element forms ions with charge of 2-. Explain (2mks)

(b) What is the nature of the oxide formed by C. (1mk)

(c) How does the reactivity of H compare with that of E. Explain? (2mks)

(d)Write down a balanced equation between B and Chlorine. (1mk)

(e) Explain how the atomic radii of F and G compare. (1mk)

(f) If the oxides of F and D are separately dissolved in water, state and explain the effects of their aqueous solutions on litmus. (3mks)

5. (a) The grid below show part of the periodic table.(The letter do not represent the actual symbols).Use it to answer the questions that follow.

T								**Q**
				S		**R**	**K**	
A	**J**		**Y**		**U**		**L**	
W							**M**	**B**
	C						**N**	
P								

(i)Select the most reactive non-metal. (1mk)

(ii)Select an element that forms a divalent cation. (1mk)

(iii)Element **Z** has atomic number 14.Show its position in the grid. (1mk)

(iv)How do the atomic radii of **U** and **J** compare? (2mks)

(v)How do electrical conductivity of **A** and **Y** compare? (2mks)

(vi)How does the boiling point of elements **K, L** and **M** vary? Explain (2mks

(b) 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 symbols of the elements.

Element	**Electron arrangement**	**Atomic radius**	**Ionic radius**
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. (2mks)

(b) Which element is a non-metal? Explain. (1mk)

(c) Which one of the elements is the strongest reducing agent. (1mk)

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

							A
			B				

	C		**D**			**E**		
	F							

(i) What name is given to the group of elements to which C and F belong? (1mk)

(ii) Which letter represents the element that is the least reactive? (1mk)

(iii) What type of bond is formed when B and E react? Explain (2mks)

(iv)Write formula of the compound formed where elements D and oxygen gas react. (1mk)

(v) On the grid indicate the a tick (√) the position of element G which is in the third period of the periodic table and forms G^{3-} ions. (1mk)

(b) Study the information in the table below and answer the questions that follow. (The letter do not represents the actual symbols of the substance).

Substance	Melting point ^{o}C	Boiling point ^{o}C	Solubility in water	Density at room. Temp/g/cm^3
H	-117	78.5	Very soluble	0.8
J	-78	-33	Very soluble	$0.77x\ 1^{-3}$
K	-23	77	Insoluble	1.6
L	- 219	-183	Slightly Soluable	1.33×10^{-3}

I.(i) Which substance would dissolve in water and could be separated from the solution by fractional distillation. (1mk)

(ii) Which substances is a liquid at room temperature and when mixed with water two layers would be formed? (1mk)

II. Which letter represents a substance that is a gas at room temperature and which can be collected ;

(i) Over water? (1mk)

(ii) By downward displacement of air? Density of air at room temperature = 1.29×10^{-3} g/C

(1mk)

UPGRADE CHEMISTRY

FORM 2

INTRODUCTION TO SALTS

1.(a) A salt is an ionic compound formed when the cation from a base combine with the anion derived from an acid.
A salt is therefore formed when the hydrogen ions in an acid are replaced wholly/fully or partially/partly ,directly or indirectly by a metal or ammonium radical.

(b) The number of ionizable/replaceable hydrogen in an acid is called **basicity** of an acid. Some acids are therefore:

(i)**mono**basic acids generally denoted **H**X e.g.

HCl, HNO_3,HCOOH,CH3COOH.

(ii)**di**basic acids ; generally denoted H_2X e.g.

H_2SO_4, H_2SO_3, H_2CO_3,HOOCOOH.

(iii)**tri**basic acids ; generally denoted H_3X e.g. H_3PO_4.

(c) Some salts are **normal** salts while other are **acid** salts.

(i)A normal salt is formed when all the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.

(ii)An acid salt is formed when part/portion the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.

Table showing normal and acid salts derived from common acids

Acid name	**Chemical formula**	**Basicity**	**Normal salt**	**Acid salt**
Hydrochloric acid	HCl	Monobasic	Chloride(Cl^-)	None
Nitric(V)acid	HNO_3	Monobasic	Nitrate(V)(NO_3^-)	None
Nitric(III)acid	HNO_2	Monobasic	Nitrate(III)(NO_2^-)	None
Sulphuric(VI)acid	H_2SO_4	Dibasic	Sulphate(VI) (SO_4^{2-})	Hydrogen sulphate(VI) (HSO_4^-)
Sulphuric(IV)acid	H_2SO_3	Dibasic	Sulphate(IV) (SO_3^{2-})	Hydrogen sulphate(IV) (HSO_3^-)
Carbonic(IV)acid	H_2CO_3	Dibasic	Carbonate(IV)(CO_3^{2-})	Hydrogen carbonate(IV) (HCO_3^-)
Phosphoric(V) acid	H_3PO_4	Tribasic	Phosphate(V)(PO_4^{3-})	Dihydrogen phosphate(V) ($H_2PO_4^{2-}$) Hydrogen diphosphate(V) ($HP_2O_4^{2-}$)

The table below show shows some examples of salts.

Base/alkali	Cation	Acid	Anion	Salt	Chemical name of salts
NaOH	Na^+	HCl	Cl^-	NaCl	Sodium(I)chloride

$Mg(OH)_2$	Mg^{2+}	H_2SO_4	SO_4^{2-}	$MgSO_4$ $Mg(HSO_4)_2$	Magnesium sulphate(VI) Magnesium hydrogen sulphate(VI)
$Pb(OH)_2$	Pb^{2+}	HNO_3	NO_3^-	$Pb(NO_3)_2$	Lead(II)nitrate(V)
$Ba(OH)_2$	Ba^{2+}	HNO_3	NO_3^-	$Ba(NO_3)_2$	Barium(II)nitrate(V)
$Ca(OH)_2$	Ba^{2+}	H_2SO_4	SO_4^{2-}	$MgSO_4$	Calcium sulphate(VI)
NH_4OH	NH_4^+	H_3PO_4	PO_4^{3-}	$(NH_4)_3PO_4$ $(NH_4)_2HPO_4$ $NH_4H_2PO_4$	Ammonium phosphate(V) Diammonium phosphate(V) Ammonium diphosphate(V)
KOH	K^+	H_3PO_4	PO_4^{3-}	K_3PO_4	Potassium phosphate(V)
$Al(OH)_3$	Al^{3+}	H_2SO_4	SO_4^{2-}	$Al_2(SO_4)_2$	Aluminium(III)sulphate(VI)
$Fe(OH)_2$	Fe^{2+}	H_2SO_4	SO_4^{2-}	$FeSO_4$	Iron(II)sulphate(VI)
$Fe(OH)_3$	Fe^{3+}	H_2SO_4	SO_4^{2-}	$Fe_2(SO_4)_2$	Iron(III)sulphate(VI)

(d) Some salts undergo **hygroscopy**, **deliquescence** and **efflorescence**.

(i) Hygroscopic salts /compounds are those that absorb water from the atmosphere but do not form a solution.

Some salts which are hygroscopic include anhydrous copper(II)sulphate(VI), anhydrous cobalt(II)chloride, potassium nitrate(V) common table salt.

(ii)Deliquescent salts /compounds are those that absorb water from the atmosphere and form a solution.

Some salts which are deliquescent include: Sodium nitrate(V),Calcium chloride, Sodium hydroxide, Iron(II)chloride, Magnesium chloride.

(iii)Efflorescent salts/compounds are those that lose their water of crystallization to the atmosphere.

Some salts which effloresces include: sodium carbonate decahydrate, Iron(II)sulphate(VI)heptahydrate, sodium sulphate (VI)decahydrate.

(e)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(III)sulphate(VI)hexahydrate	$Al_2 (SO_4)_3.6H_2O$
Zinc(II)sulphate(VI)heptahydrate	$ZnSO_4.7H_2O$
Iron(II)sulphate(VI)heptahydrate	$FeSO_4.7H_2O$
Calcium(II)sulphate(VI)heptahydrate	$CaSO_4.7H_2O$
Magnesium(II)sulphate(VI)heptahydrate	$MgSO_4.7H_2O$
Sodium sulphate(VI)decahydrate	$Na_2SO_4.10H_2O$
Sodium carbonate(IV)decahydrate	$Na_2CO_3.10H_2O$
Potassium carbonate(IV)decahydrate	$K_2CO_3.10H_2O$
Potassium sulphate(VI)decahydrate	$K_2SO_4.10H_2O$

(f)Some salts exist as a simple salt while some as complex salts. Below are some complex salts.

Table of some complex salts

Name of complex salt	Chemical formula	Colour of the complex salt
Tetraamminecopper(II)sulphate(VI)	$Cu(NH_3)_4 SO_4 H_2O$	Royal/deep blue solution
Tetraamminezinc(II)nitrate(V)	$Zn(NH_3)_4 (NO_3)_2$	Colourless solution
Tetraamminecopper(II) nitrate(V)	$Cu(NH_3)_4 (NO_3)_2$	Royal/deep blue solution
Tetraamminezinc(II)sulphate(VI)	$Zn(NH_3)_4 SO_4$	Colourless solution

(g)Some salts exist as two salts in one. They are called **double salts**.

Table of some double salts

Name of double salts	Chemical formula
Trona(sodium sesquicarbonate)	$Na_2CO_3 NaHCO_3.2H_2O$
Ammonium iron(II)sulphate(VI)	$FeSO_4(NH_4)_2SO_4.2H_2O$
Ammonium aluminium(III)sulphate(VI)	$Al_2(SO_4)_3(NH_4)_2SO_4.H_2O$

(h)Some salts dissolve in water to form a **solution**. They are said to be soluble. Others do not dissolve in water. They form a **suspension/precipitate** in water.

Table of solubility of salts

Soluble salts	Insoluble salts
All nitrate(V)salts	
All sulphate(VI)/SO_4^{2-} salts **except** →	Barium(II) sulphate(VI)/$BaSO_4$ Calcium(II) sulphate(VI)/$CaSO_4$ Lead(II) sulphate(VI)/$PbSO_4$
All sulphate(IV)/SO_3^{2-} salts **except** →	Barium(II) sulphate(IV)/$BaSO_3$ Calcium(II) sulphate(IV)/$CaSO_3$ Lead(II) sulphate(IV)/$PbSO_3$
All chlorides/Cl^- **except** →	Silver chloride/AgCl Lead(II)chloride/$PbCl_2$(dissolves in hot water)

All phosphate(V)/PO_4^{3-}	
All sodium,potassium and ammonium salts	
All hydrogen carbonates/HCO_3^-	
All hydrogen sulphate(VI)/ HSO_4^-	
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

13 Salts can be prepared in a school laboratory by a method that uses its solubility in water.
(a) Soluble salts may be prepared by using any of the following methods:

(i)Direct displacement/reaction of a metal with an acid.

By reacting a metal higher in the reactivity series than hydrogen with a dilute acid,a salt is formed and hydrogen gas is evolved.
Excess of the metal must be used to ensure all the acid has reacted.
When effervescence/bubbling /fizzing has stopped ,excess metal is filtered.
The filtrate is heated to concentrate then allowed to crystallize.
Washing with distilled water then drying between filter papers produces a sample crystal of the salt. i.e.

M(s) + H_2X -> MX(aq) + H_2(g)

Examples

Mg(s) + H_2SO_4(aq) -> $MgSO_4$ (aq) + H_2(g)
Zn(s) + H_2SO_4(aq) -> $ZnSO_4$ (aq) + H_2(g)
Pb(s) + 2HNO_3(aq) -> $Pb(NO_3)_2$(aq) + H_2(g)
Ca(s) + 2HNO_3(aq) -> $Ca(NO_3)_2$(aq) + H_2(g)
Mg(s) + 2HNO_3(aq) -> $Mg(NO_3)_2$(aq) + H_2(g)
Mg(s) + 2HCl(aq) -> $MgCl_2$(aq) + H_2(g)
Zn(s) + 2HCl(aq) -> $ZnCl_2$(aq) + H_2(g)

(ii)Reaction of an insoluble base with an acid

By adding an insoluble base (oxide/hydroxide)to a dilute acid until **no more dissolves**, in the acid,a salt and water are formed. Excess of the base is filtered off. The filtrate is heated to concentrate ,allowed to crystallize then washed with distilled water before drying between filter papers e.g.

PbO(s) + 2HNO_3(aq) -> $Pb(NO_3)_2$(aq) + H_2O (l)
$Pb(OH)_2$(s) + 2HNO_3(aq) -> $Pb(NO_3)_2$(aq) + 2H_2O (l)
CaO (s) + 2HNO_3(aq) -> $Ca(NO_3)_2$(aq) + H_2O (l)
MgO (s) + 2HNO_3(aq) -> $Mg(NO_3)_2$(aq) + H_2O (l)
MgO (s) + 2HCl(aq) -> $MgCl_2$(aq) + H_2O (l)
ZnO (s) + 2HCl(aq) -> $ZnCl_2$(aq) + H_2O (l)
$Zn(OH)_2$(s) + 2HNO_3(aq) -> $Zn(NO_3)_2$(aq) + 2H_2O (l)
CuO (s) + 2HCl(aq) -> $CuCl_2$(aq) + H_2O (l)
CuO (s) + H_2SO_4(aq) -> $CuSO_4$(aq) + H_2O (l)

$Ag_2O(s) + 2HNO_3(aq) \rightarrow 2AgNO_3(aq) + H_2O(l)$
$Na_2O(s) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l)$

(iii)reaction of insoluble /soluble carbonate /hydrogen carbonate with an acid.

By adding an excess of a soluble /insoluble carbonate or hydrogen carbonate to adilute acid, effervescence /fizzing/bubbling out of carbon(IV)oxide gas shows the reaction is taking place. When effervescence /fizzing/bubbling out of the gas is over, excess of the insoluble carbonate is filtered off. The filtrate is heated to concentrate ,allowed to crystallize then washed with distilled water before drying between filter paper papers e.g.

$PbCO_3(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
$ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
$MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$
$Cu\,CO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$
$Ag_2CO_3(s) + 2HNO_3(aq) \rightarrow 2AgNO_3(aq) + H_2O(l) + CO_2(g)$
$Na_2CO_3(s) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l) + CO_2(g)$
$K_2CO_3(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2O(l) + CO_2(g)$
$NaHCO_3(s) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l) + CO_2(g)$
$KHCO_3(s) + HCl(aq) \rightarrow KCl(aq) + H_2O(l) + CO_2(g)$

(iv)neutralization/reaction of soluble base/alkali with dilute acid

By adding an acid to a burette into a known volume of an alkali with 2-3 drops of an indicator, the colour of the indicator changes when the acid has completely reacted with an alkali at the **end point.** The procedure is then repeated without the indicator .The solution mixture is then heated to concentrate , allowed to crystallize ,washed with distilled water before drying with filter papers. e.g.

$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$
$KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$
$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
$2\,NH_4OH(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + 2H_2O(l)$
$NH_4OH(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq) + H_2O(l)$

(iv)direct synthesis/combination.

When a metal **burn** in a gas jar containing a non metal , the two directly combine to form a salt. e.g.

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$
$Mg(s) + Cl_2(g) \rightarrow Mg\,Cl_2(s)$
$Ca(s) + Cl_2(g) \rightarrow Ca\,Cl_2(s)$

Some salts once formed undergo **sublimation** and **hydrolysis**. Care should be taken to avoid water/moisture into the reaction flask during their preparation.Such salts include aluminium(III)chloride($AlCl_3$) and iron (III)chloride($FeCl_3$)

1. Heated aluminium foil reacts with chlorine to form aluminium(III)chloride that sublimes away from the source of heating then deposited as solid again

$2Al(s) \quad + \quad 3Cl_2(g) \quad \rightarrow \quad 2AlCl_3\ (\mathbf{s/g})$

Once formed aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

$AlCl_3(s) + \quad 3H_2O(g) \quad \rightarrow \quad Al(OH)_3\ (\mathbf{aq}) + 3HCl(g)$

2. Heated iron filings reacts with chlorine to form iron(**III**)chloride that sublimes away from the source of heating then deposited as solid again

$2Fe(s) \quad + \quad 3Cl_2(g) \quad \rightarrow \quad 2FeCl_3\ (\mathbf{s/g})$

Once formed , aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

$FeCl_3(s) + \quad 3H_2O(g) \quad \rightarrow \quad Fe(OH)_3\ (\mathbf{aq}) + 3HCl(g)$

(b)**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.

$\mathbf{Cu}SO_4(aq) \quad + \quad Na_2\mathbf{CO_3}\ (aq) \quad \rightarrow \quad \mathbf{CuCO_3}\ (s) \ + \ Na_2SO_4(aq)$

$\mathbf{Ba}Cl_2(aq) \quad + \quad K_2\mathbf{SO_4}\ (aq) \quad \rightarrow \quad \mathbf{BaSO_4}\ (s) \ + \ 2KCl\ (aq)$

$\mathbf{Pb(}NO_3)_2(aq) \quad + \quad K_2\mathbf{SO_4}\ (aq) \quad \rightarrow \quad \mathbf{PbSO_4}\ (s) \ + \ 2KNO_3\ (aq)$

$2\mathbf{Ag}NO_3(aq) \quad + \quad Mg\mathbf{Cl_2}\ (aq) \quad \rightarrow \quad 2\mathbf{AgCl}(s) \ + \ Mg(NO_3)_2\ (aq)$

$\mathbf{Pb}(NO_3)_2(aq) \quad + \quad (NH_4)_2\mathbf{SO_4}\ (aq) \quad \rightarrow \quad \mathbf{PbSO_4}\ (s) \ + \ 2NH_4NO_3(aq)$

$\mathbf{Ba}Cl_2(aq) \quad + \quad K_2\mathbf{SO_3}\ (aq) \quad \rightarrow \quad \mathbf{BaSO_3}\ (s) \ + \ 2KCl\ (aq)$

14. Salts may lose their water of crystallization , decompose ,melt or sublime on heating on a Bunsen burner flame.

The following shows the behavior of some salts on heating gently /or strongly in a laboratory school burner:

(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 **sublimes** on **gentle** heating. It **dissociate** into the constituent **ammonia** and **hydrogen chloride** gases on strong heating.

$$NH_4Cl(s) \underset{\text{(sublimation)}}{\rightleftharpoons} NH_4Cl(g) \underset{\text{(dissociation)}}{\rightleftharpoons} NH_3(g) + HCl(g)$$

(b)effect of heat on nitrate(V)

(i) Potassium nitrate(V)/KNO_3 and sodium nitrate(V)/$NaNO_3$ decompose on heating to form Potassium nitrate(III)/KNO_2 and sodium nitrate(III)/$NaNO_2$ and producing Oxygen gas in each case.

$2KNO_3$ (s) -> $2KNO_2$(s) + O_2(g)
$2NaNO_3$ (s) -> $2NaNO_2$(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) -> 2CaO(s) + $4NO_2$(g) + O_2(g)
$2Mg(NO_3)_2$(s) -> 2MgO(s) + $4NO_2$(g) + O_2(g)
$2Zn(NO_3)_2$(s) -> 2ZnO(s) + $4NO_2$(g) + O_2(g)
$2Pb(NO_3)_2$(s) -> 2PbO(s) + $4NO_2$(g) + O_2(g)
$2Cu(NO_3)_2$(s) -> 2CuO(s) + $4NO_2$(g) + O_2(g)
$2Fe(NO_3)_2$(s) -> 2FeO(s) + $4NO_2$(g) + O_2(g)

(iii)Silver(I)nitrate(V) and mercury(II) nitrate(V) 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) -> 2Ag (s) + $2NO_2$(g) + O_2(g)
$2Hg(NO_3)_2$ (s) -> 2Hg (s) + $4NO_2$(g) + O_2(g)

(iv)Ammonium nitrate(V) and Ammonium nitrate(III) 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) -> N_2O (g) + H_2O(l)
NH_4NO_2(s) -> N_2 (g) + H_2O(l)

(c) effect of heat on nitrate(V)

Only Iron(II)sulphate(VI), Iron(III)sulphate(VI) and copper(II)sulphate(VI) decompose on heating. They form the **oxide**, and produce highly acidic fumes of acidic **sulphur(IV)oxide** gas.

$2FeSO_4$ (s) -> Fe_2O_3(s) + SO_3(g) + SO_2(g)
$Fe_2(SO_4)_3$(s) -> Fe_2O_3(s) + SO_3(g)
$CuSO_4$ (s) -> CuO(s) + SO_3(g)

(d) effect of heat on carbonates(IV) and hydrogen carbonate(IV).

(i)Sodium carbonate(IV)and potassium carbonate(IV)**do not decompose** on heating.

(ii)Heavy metal nitrate(IV)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) -> CuO(s) + CO_2(g)

$CaCO_3$ (s) -> CaO(s) + CO_2(g)
$PbCO_3$ (s) -> PbO(s) + CO_2(g)
$FeCO_3$ (s) -> FeO(s) + CO_2(g)
$ZnCO_3$ (s) -> ZnO(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.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$$

$$2KHCO_3(s) \rightarrow K_2CO_3(s) + CO_2(g) + H_2O(l)$$

(iii) 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.

$$Ca(HCO_3)_2(\mathbf{aq}) \rightarrow CaCO_3(s) + CO_2(g) + H_2O(l)$$

$$Mg(HCO_3)_2(\mathbf{aq}) \rightarrow MgCO_3(s) + CO_2(g) + H_2O(l)$$

INTRODUCTION TO ELECTROLYSIS (ELECTROLYTIC CELL)

1.Electrolysis is defined simply as the **decomposition** of a **compound** by an electric current/**electricity**.
A compound that is decomposed by an electric current is called an electrolyte. Some electrolytes are **weak** while others are **strong**.

2.Strong electrolytes are those that are fully ionized/dissociated into (many) ions. Common strong electrolytes include:

(i)all **mineral** acids
(ii)all strong **alkalis**/sodium hydroxide/potassium hydroxide.
(iii)all soluble **salts**

3.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**

4. A compound that is **not** decomposed by an electric current is called non-electrolyte.

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)

5. An electrolytes in **solid** state have **fused** /joined ions and therefore do **not** conduct electricity but the **ions** (cations and anions) are **free** and **mobile** in **molten** and **aqueous** (solution, dissolved in water) state.

6.During electrolysis, the free ions are attracted to the **electrodes**.

An electrode is a rod through which current enter and leave the electrolyte during electrolysis.

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).

7.The **positive** electrode is called **Anode**.The anode is the electrode through which **current enter** the electrolyte/**electrons leave** the electrolyte

8.The **negative** electrode is called **Cathode**. The cathode is the electrode through which **current leave** the electrolyte / **electrons enter** the electrolyte

9. During the electrolysis, free **anions** are attracted to the **anode** where they **lose** /**donate** electrons to form **neutral** atoms/molecules. i.e.

M(l) -> M^+(l) + e (for cations from molten electrolytes)

M(s) -> M^+(aq) + e (for cations from electrolytes in aqueous state / solution / dissolved in water)

The neutral atoms /molecules form the **products** of electrolysis at the anode. This is called **discharge** at anode

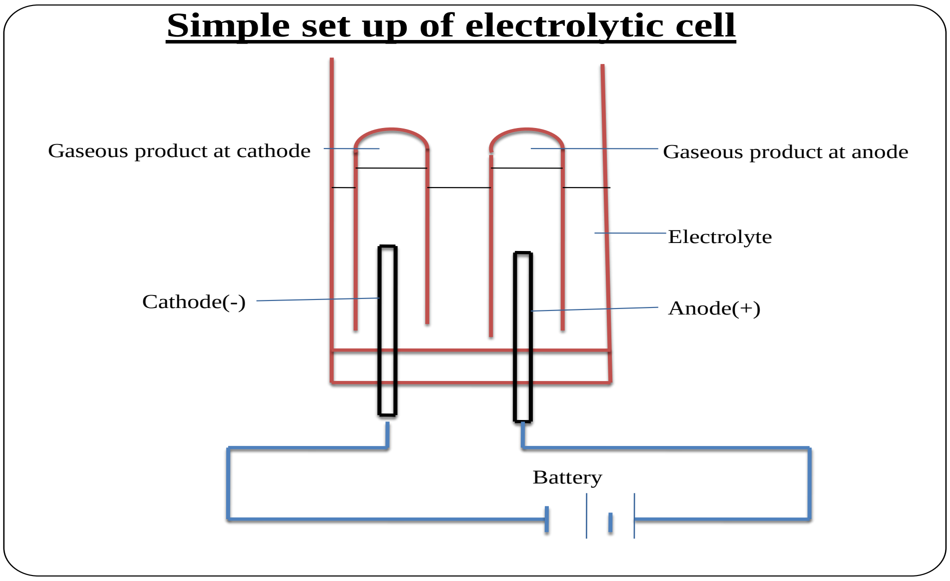
10. During electrolysis, free **cations** are attracted to the **cathode** where they **gain /accept/acquire** electrons to form **neutral** atoms/molecules.

X^+ (aq) + 2e -> X(s) (for cations from electrolytes in aqueous state / solution / dissolved in water)

$2X^+$ (l) + 2e -> X (l) (for cations from molten electrolytes)

The neutral atoms /molecules form the **products** of electrolysis at the cathode. This is called **discharge** at cathode.

11. The below set up shows an electrolytic cell.



Simple set up of electrolytic cell

12. For a compound /salt containing only two ion/binary salt the products of electrolysis in an electrolytic cell can be determined as in the below examples:

a)<u>To determine the products of electrolysis of molten Lead(II)chloride</u>

(i)Decomposition of electrolyte into free ions;

$PbCl_2$ (l) -> Pb^{2+}(l) + $2Cl^-$(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

Pb^{2+}(l) + 2e -> Pb (l)

(Cation / Pb^{2+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$2Cl^-$(l) -> Cl_2 (g) + 2e

(Anion / Cl^- donate/lose electrons to form free **atom** then a gas **molecule)** (iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid lead metal.

II.At the anode pale green chlorine gas.

b)To determine the products of electrolysis of molten Zinc bromide

(i)Decomposition of electrolyte into free ions;

$ZnBr_2 (l) \rightarrow Zn^{2+}(l) + 2Br^-(l)$

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

$Zn^{2+}(l) + 2e \rightarrow Zn(l)$

(Cation / Zn^{2+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$2Br^-(l) \rightarrow Br_2 (g) + 2e$

(Anion / Br^- donate/lose electrons to form free **atom** then a liquid **molecule** which change to **gas** on heating)

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid Zinc metal.

II.At the anode **red** bromine **liquid** / **red/brown** bromine **gas.**

c)To determine the products of electrolysis of molten sodium chloride

(i)Decomposition of electrolyte into free ions;

$NaCl (l) \rightarrow Na^+(l) + Cl^-(l)$

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

$2Na^+(l) + 2e \rightarrow Na (l)$

(Cation / Na^+ gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$2Cl^-(l) \rightarrow Cl_2 (g) + 2e$

(Anion / Cl^- donate/lose electrons to form free **atom** then a gas **molecule)**

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid sodium metal.

II.At the anode pale green chlorine gas.

d)To determine the products of electrolysis of molten Aluminium (III)oxide

(i)Decomposition of electrolyte into free ions;

$Al_2O_3 (l) \rightarrow 2Al^{3+}(l) + 3O^{2-}(l)$

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

$4Al^{3+}(l) + 12e \rightarrow 4Al(l)$

(Cation / Al^{3+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$6O^{2-}(l) \rightarrow 3O_2(g) + 12e$

(Anion /$6O^{2-}$ donate/lose 12 electrons to form free **atom** then three gas **molecule)**

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid aluminium metal.
II.At the anode colourless gas that relights/rekindles glowing splint.

13. In industries electrolysis has the following uses/applications:

(a)Extraction of reactive metals from their ores.

Potassium, sodium ,magnesium, and aluminium are extracted from their ores using electrolytic methods.

(b)Purifying copper after exraction from copper pyrites ores.

Copper obtained from copper pyrites ores is not pure. After extraction, the copper is refined by electrolysing copper(II)sulphate(VI) solution using the **impure** copper as **anode** and a thin strip of **pure** copper as **cathode**. Electrode ionization take place there:

(i)At the cathode; $\mathbf{Cu^{2+}(aq) + 2e \rightarrow Cu(s)}$ (Pure copper deposits on the strip

(ii)At the anode; $\mathbf{Cu(s) \rightarrow Cu^{2+}(aq) + 2e}$ (impure copper erodes/dissolves)

(c)Electroplating

The label EPNS(**E**lectro **P**lated **N**ickel **S**ilver) on some steel/metallic utensils mean they are plated/coated with silver and/or Nickel to **improve** their **appearance**(**add** their **aesthetic** value)and **prevent**/slow **corrosion**(**rusting** of iron). Electroplating is the process of coating a metal with another metal using an electric current. During electroplating, the **cathode** is made of the metal to be **coated**/impure.

Example:

During the electroplating of a spoon with silver

(i)the spoon/impure is placed as the cathode(negative terminal of battery)
(ii)the pure silver is placed as the anode(positive terminal of battery)
(iii)the pure silver erodes/ionizes/dissociates to release electrons:

$\mathbf{Ag(s) \rightarrow Ag^{+}(aq) + e}$ (impure silver erodes/dissolves)

(iv) silver ($\mathbf{Ag^{+}}$)ions from electrolyte gain electrons to form pure silver deposits / coat /cover the spoon/impure

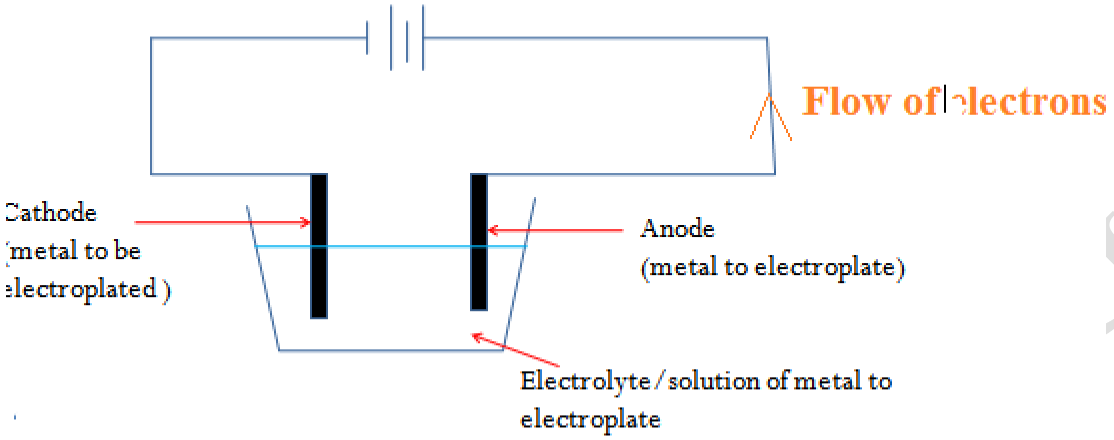
$\mathbf{Ag^{+}(aq) + e \rightarrow Ag(s)}$ (pure silver deposits /coat/cover on spoon)

$Ag(s) \rightarrow Ag^+(aq) + e$ (pure silver erodes/dissolves)

(iv) silver (**Ag^+**)ions from electrolyte gain electrons to form pure silver deposits / coat /cover the spoon/impure

$Ag^+(aq) + e \rightarrow Ag(s)$ (pure silver deposits /coat/cover on spoon)

<u>Electrolytic set up during electroplating</u>



A: CARBON

Carbon is an element in Group IV(Group 4)of the Periodic table .It has atomic number 6 and electronic configuration 2:4 and thus has four valence electrons(tetravalent).It does not easily ionize but forms strong covalent bonds with other elements including itself.

(a)Occurrence

Carbon mainly naturally occurs as:

(i)allotropes of carbon i.e graphite, diamond and fullerenes.

(ii)amorphous carbon in coal, peat ,charcoal and coke.

(iii)carbon(IV)oxide gas accounting 0.03% by volume of normal air in the atmosphere.

(b)Allotropes of Carbon

Carbon naturally occur in two main crystalline allotropic forms, carbon-graphite and carbon-diamond

Carbon-diamond	Carbon-graphite
Shiny crystalline solid	Black/dull crystalline solid
Has a very high melting/boiling point because it has a very closely packed giant tetrahedral structure joined by strong covalent bonds	Has a high melting/boiling point because it has a very closely packed giant hexagonal planar structure joined by strong covalent bonds
Has very high density(Hardest known natural substance)	Soft
Abrassive	Slippery
Poor electrical conductor because it has no free delocalized electrons	Good electrical conductor because it has free 4th valency delocalized electrons
Is used in making Jewels, drilling and cutting metals	Used in making Lead-pencils,electrodes in batteries and as a lubricant
Has giant tetrahedral structure	Has giant hexagonal planar structure

c)Properties of Carbon

(i)Physical properties of carbon

Carbon occur widely and naturally as a black solid

It is insoluble in water but soluble in carbon disulphide and organic solvents.

It is a poor electrical and thermal conductor.

(ii)Chemical properties of carbon

I. Burning

Experiment

Introduce a small piece of charcoal on a Bunsen flame then lower it into a gas jar containing Oxygen gas. Put three drops of water. Swirl. Test the solution with blue and red litmus papers.

Observation

-Carbon chars then burns with a blue flame

-Colourless and odourless gas produced

-Solution formed turn blue litmus paper faint red. Red litmus paper remains red.

Explanation

Carbon burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon burns in limited supply of air with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon (IV) oxide gas dissolve in water to form weak acidic solution of Carbonic (IV)acid.

Chemical Equation

$C(s) + O_2(g) \rightarrow CO_2(g)$ (in excess air)

$2C(s) + O_2(g) \rightarrow 2CO(g)$ (in limited air)

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$ (very weak acid)

II. Reducing agent

Experiment

Mix thoroughly equal amounts of powdered charcoal and copper (II)oxide into a crucible. Heat strongly.

Observation

Colour change from black to brown

Explanation

Carbon is a reducing agent. For ages it has been used to reducing metal oxide ores to metal, itself oxidized to carbon(IV)oxide gas. Carbon reduces black copper(II)oxide to brown copper metal

Chemical Equation

$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$

(black) (brown)

$2PbO(s) + C(s) \rightarrow 2Pb(s) + CO_2(g)$

(brown when hot/ yellow when cool) (grey)

$2ZnO(s) + C(s) \rightarrow 2Zn(s) + CO_2(g)$

(yellow when hot/ white when cool) (grey)

$Fe_2O_3(s) + 3C(s) \rightarrow 2Fe(s) + 3CO_2(g)$

(brown when hot/cool (grey)

$Fe_3O_4(s) + 4C(s) \rightarrow 3Fe(s) + 4CO_2(g)$

(brown when hot/cool (grey)

B: COMPOUNDS OF CARBON

The following are the main compounds of Carbon

(i)Carbon(IV)Oxide(CO_2)
(ii)Carbon(II)Oxide(CO)
(iii)Carbonate(IV) (CO_3^{2-})and hydrogen carbonate(IV(HCO_3^-)
(iv)Sodium carbonate(Na_2CO_3)

(i) Carbon(IV)Oxide (CO_2)

(a)Occurrence

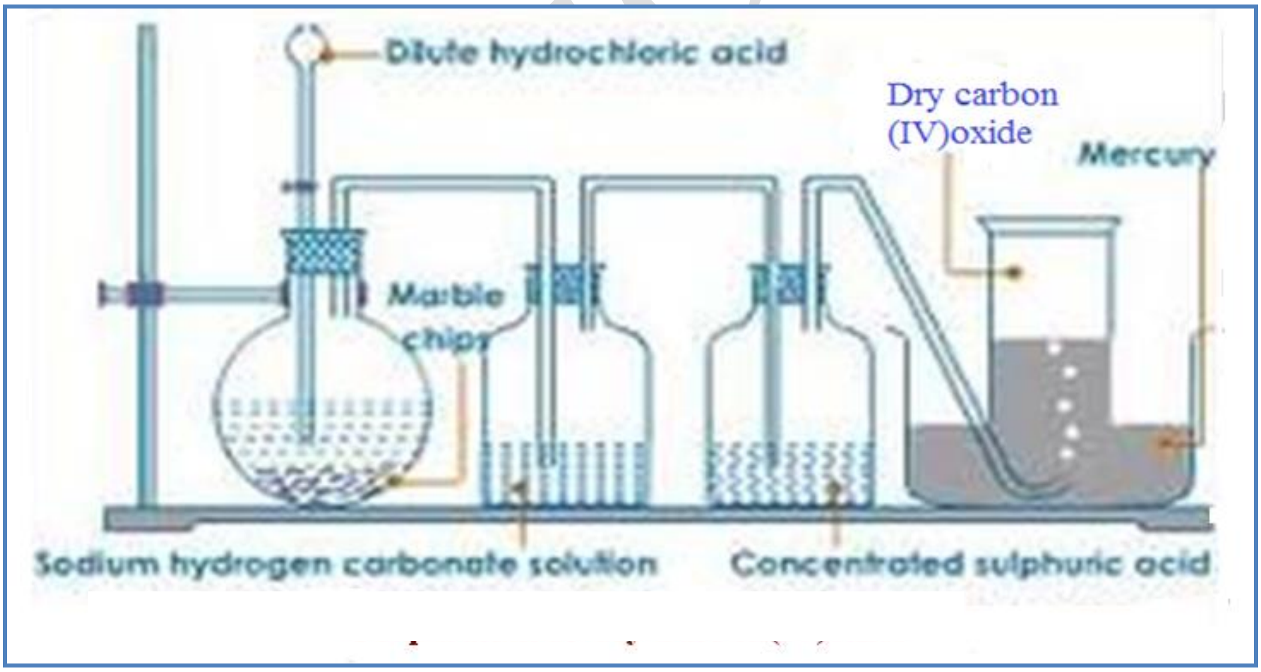
Carbon(IV)oxide is found:

-in the air /atmosphere as 0.03% by volume.

-a solid carbon(IV)oxide mineral in Esageri near Eldame Ravine and Kerita near Limuru in Kenya.

(b)School Laboratory preparation

In the school laboratory carbon(IV)oxide can be prepared in the school laboratory from the reaction of marble chips($CaCO_3$)or sodium hydrogen carbonate($NaHCO_3$) with dilute hydrochloric acid.



(c)Properties of carbon(IV)oxide gas(Questions)

1.Write the equation for the reaction for the school laboratory preparation of carbon (IV)oxide gas.

Any carbonate reacted with dilute hydrochloric acid should be able to generate carbon (IV)oxide gas.

Chemical equations

$CaCO_3(s)$ + $2HCl(aq)$ -> $CaCO_3$ (aq) + $H_2O(l)$ + CO_2 (g)
$ZnCO_3(s)$ + $2HCl(aq)$ -> $ZnCO_3$ (aq) + $H_2O(l)$ + CO_2 (g)
$MgCO_3(s)$ + $2HCl(aq)$ -> $MgCO_3$ (aq) + $H_2O(l)$ + CO_2 (g)
$CuCO_3(s)$ + $2HCl(aq)$ -> $CuCO_3$ (aq) + $H_2O(l)$ + CO_2 (g)
$NaHCO_3(s)$ + $HCl(aq)$ -> Na_2CO_3 (aq) + $H_2O(l)$ + CO_2 (g)
$KHCO_3(s)$ + $HCl(aq)$ -> K_2CO_3 (aq) + $H_2O(l)$ + CO_2 (g)

2.What method of gas collection is used in preparation of Carbon(IV)oxide gas. Explain.

Downward delivery /upward displacement of air/over mercury
Carbon(IV)oxide gas is about 1½ times denser than air.

3.What is the purpose of :

(a)water?

To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction.

(b)sodium hydrogen carbonate?

To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction and by reacting with the acid to produce more carbon (IV)oxide gas .

Chemical equation

$NaHCO_3(s)$ + $HCl(aq)$ -> Na_2CO_3 (aq) + $H_2O(l)$ + CO_2 (g)

(c)concentrated sulphuric(VI)acid?

To dry the gas/as a drying agent

4.Describe the smell of carbon(IV)oxide gas

Colourless and odourless

5. Effect on lime water.

Experiment

Bubbled carbon(IV)oxide gas into a test tube containing lime water for about three minutes

Observation

White precipitate is formed.
White precipitate dissolved when excess carbon(IV)oxide gas is bubbled .

Explanation

Carbon(IV)oxide gas reacts with lime water($Ca(OH)_2$) to form an insoluble white precipitate of calcium carbonate. Calcium carbonate reacts with more Carbon(IV) oxide gas to form soluble Calcium hydrogen carbonate.

Chemical equation

$Ca(OH)_2(aq)$ + CO_2 (g) -> $CaCO_3$ (s) + $H_2O(l)$
$CaCO_3$ (aq) + $H_2O(l)$ + CO_2 (g) -> $Ca(HCO_3)_2$ (aq)

6. Effects on burning Magnesium ribbon

Experiment

Lower a piece of burning magnesium ribbon into a gas jar containing carbon (IV)oxide gas.

Observation

The ribbon continues to burn with difficulty

White ash/solid is formed.

Black speck/solid/particles formed on the side of gas jar.

Explanation

Carbon(IV)oxide gas does not support combustion/burning.Magnesium burn to produce/release enough heat energy to decompose Carbon(IV) oxide gas to carbon and oxygen.Magnesium continues to burn in Oxygen forming white Magnesium Oxide solid/ash.Black speck/particle of carbon/charcoal residue forms on the sides of reaction flask. During the reaction Carbon(IV) oxide is reduced(Oxidizing agent)to carbon while Magnesium is Oxidized to Magnesium Oxide.

Chemical equation

$2Mg(s) + CO_2(g) \rightarrow C(s) + 2MgO(l)$

7. Dry and wet litmus papers were separately put in a gas jar containing dry carbon (IV)oxide gas. State and explain the observations made.

Observation

Blue dry litmus paper remain blue

Red dry litmus paper remain Red

Blue wet/damp/moist litmus paper turn red

Red wet/damp/moist litmus paper remain red

Explanation

Dry Carbon (IV) oxide gas is a molecular compound that does not dissociate/ionize to release $\mathbf{H^+}$ and thus has no effect on litmus papers.

Wet/damp/moist litmus papers contains water that dissolves/react with dry carbon (IV) oxide gas to form the weak solution of carbonic (IV) acid(H_2CO_3).

Carbonic (IV) acid dissociate/ionizes to a few /little free $\mathbf{H^+}$ and $\mathbf{CO_3^{2-}}$**.**

The few H^+ (aq) ions are responsible for turning blue litmus paper to faint red showing the gas is very weakly acidic.

Chemical equation

$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$

8. Explain why Carbon (IV)oxide cannot be prepared from the reaction of:

(i) marble chips with dilute sulphuric(VI)acid.

Explanation

Reaction forms insoluble calcium sulphate(VI)that cover/coat unreacted marble chips stopping further reaction

Chemical equation

$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(\mathbf{s}) + H_2O(l) + CO_2(g)$

$PbCO_3(s) + H_2SO_4(aq) \rightarrow PbSO_4(\mathbf{s}) + H_2O(l) + CO_2(g)$

$BaCO_3(s) + H_2SO_4(aq) \rightarrow BaSO_4(\mathbf{s}) + H_2O(l) + CO_2(g)$

(ii) Lead(II)carbonate with dilute Hydrochloric acid.

Reaction forms insoluble Lead(II)Chloride that cover/coat unreacted Lead(II) carbonate stopping further reaction unless the reaction mixture is heated. Lead(II)Chloride is soluble in hot water.

Chemical equation

$PbCO_3(s)$ + $2HCl(aq)$ -> $PbCl_2(s)$ + $H_2O(l)$ + $CO_2(g)$

9. Describe the test for the presence of Carbon (IV)oxide.

Using burning splint

Lower a burning splint into a gas jar suspected to contain Carbon (IV)oxide gas.The burning splint is extinguished.

Using Lime water.

Bubble the gas suspected to be Carbon (IV)oxide gas.A white precipitate that dissolve in excess bubbling is formed.

Chemical equation

$Ca(OH)_2(aq)$ + $CO_2(g)$ -> $CaCO_3(s)$ + $H_2O(l)$

$CaCO_3(aq)$ + $H_2O(l)$ + $CO_2(g)$ -> $Ca(HCO_3)_2(aq)$

10.State three main uses of Carbon (IV)oxide gas

(i)In the Solvay process for the manufacture of soda ash/sodium carbonate

(ii)In preservation of aerated drinks

(iii)As fire extinguisher because it does not support combustion and is denser than air.

(iv)In manufacture of Baking powder.

(ii) Carbon(II)Oxide (CO)

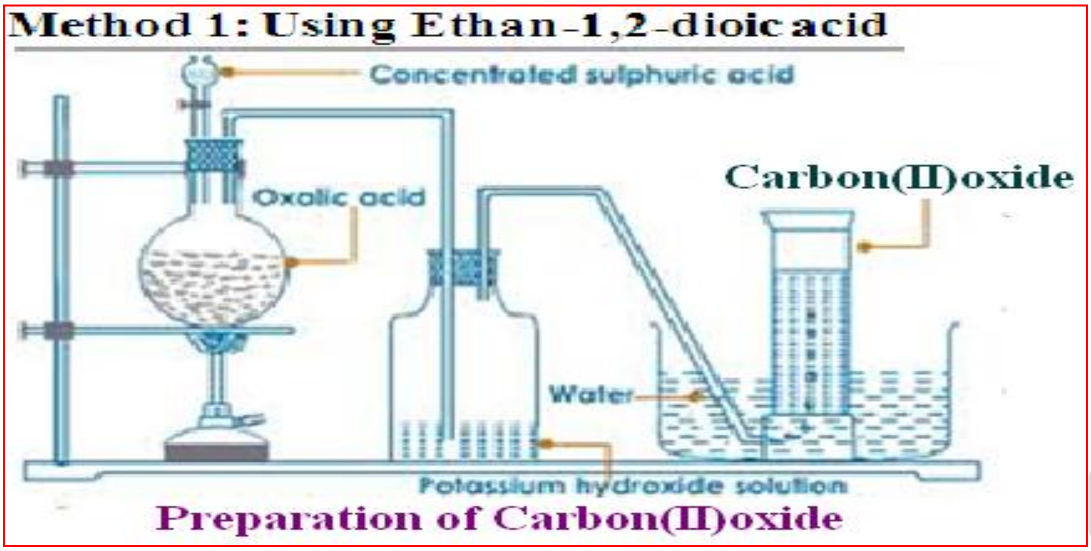
(a)Occurrence

Carbon(II)oxide is found is found from incomplete combustion of fuels like petrol charcoal, liquefied Petroleum Gas/LPG.

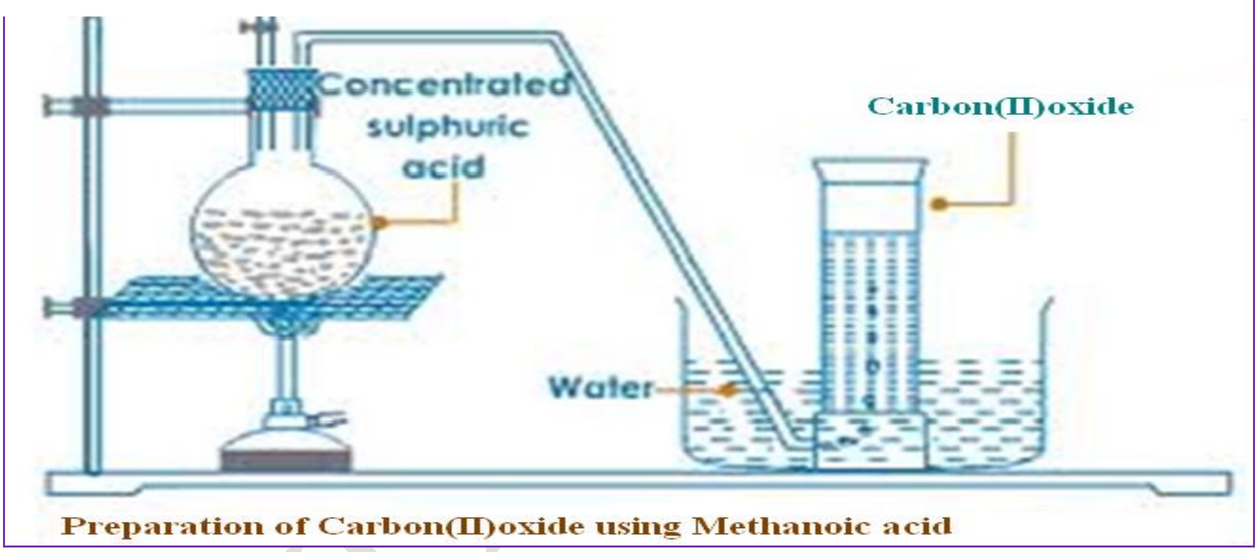
(b)School Laboratory preparation

In the school laboratory carbon(II)oxide can be prepared from dehydration of methanoic acid/Formic acid(HCOOH) or Ethan-1,2-dioic acid/Oxalic acid(HOOCCOOH) using concentrated sulphuric(VI) acid. Heating is necessary.

METHOD 1:Preparation of Carbon (IV)Oxide from dehydration of Oxalic/ethan-1,2-dioic acid



METHOD 2:Preparation of Carbon (IV)Oxide from dehydration of Formic/Methanoic acid



(c)Properties of Carbon (II)Oxide(Questions)

1.Write the equation for the reaction for the preparation of carbon(II)oxide using;

(i)Method 1;

Chemical equation

$HOOCCOOH(s)$ –Conc.H_2SO_4--> $CO(g)$ + $CO_2(g)$ + $H_2O(l)$

$H_2C_2O_4(s)$ –Conc.H_2SO_4--> $CO(g)$ + $CO_2(g)$ + $H_2O(l)$

(ii)Method 2;

Chemical equation

$HCOOH(s)$ –Conc.H_2SO_4--> $CO(g)$ + $H_2O(l)$

$H_2CO_2(s)$ –Conc.H_2SO_4--> $CO(g)$ + $H_2O(l)$

2.What method of gas collection is used during the preparation of carbon (II) oxide.

Over water because the gas is insoluble in water.

Downward delivery because the gas is 1 ½ times denser than air .

3.What is the purpose of :

(i) Potassium hydroxide/sodium hydroxide in Method 1

To absorb/ remove carbon (II) oxide produced during the reaction.

$2KOH (aq) + CO_2 (g) \rightarrow K_2CO_3 (s) + H_2O(l)$

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

(ii) Concentrated sulphuric(VI)acid in Method 1 and 2.

Dehydrating agent –removes the element of water (Hydrogen and Oxygen in ratio 2:1) present in both methanoic and ethan-1,2-dioic acid.

4. Describe the smell of carbon(II)oxide.

Colourless and odourless.

5. State and explain the observation made when carbon(IV)oxide is bubbled in lime water for a long time.

No white precipitate is formed.

6. Dry and wet/moist/damp litmus papers were separately put in a gas jar containing dry carbon(IV)oxide gas. State and explain the observations made.

Observation

-blue dry litmus paper remains blue

-red dry litmus paper remains red

- wet/moist/damp blue litmus paper remains blue

- wet/moist/damp red litmus paper remains red

Explanation

Carbon(II)oxide gas is a molecular compound that does not dissociate /ionize to release H+ ions and thus has no effect on litmus papers. Carbon(II)oxide gas is therefore a **neutral** gas.

7. Carbon (II)oxide gas was ignited at the end of a generator as below.

(i)State the observations made in flame K.

Gas burns with a blue flame

(ii)Write the equation for the reaction taking place at flame K.

$2CO(g) + O_2 (g) \rightarrow 2CO_2 (g)$

8. Carbon(II)oxide is a reducing agent. Explain

Experiment

Pass carbon(II)oxide through glass tube containing copper (II)oxide. Ignite any excess poisonous carbon(II)oxide.

Observation

Colour change from black to brown. Excess carbon (II)oxide burn with a blue flame.

Explanation

Carbon is a reducing agent. It is used to reduce metal oxide ores to metal, itself oxidized to carbon(IV)oxide gas. Carbon(II)Oxide reduces black copper(II)oxide to brown copper metal

Chemical Equation

$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$
(black) (brown)

$PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$
(brown when hot/ yellow when cool) (grey)

$ZnO(s) + CO(g) \rightarrow Zn(s) + CO_2(g)$
(yellow when hot/ white when cool) (grey)

$Fe_2O_3(s) + 3CO(s) \rightarrow 2Fe(s) + 3CO_2(g)$
(brown when hot/cool (grey)

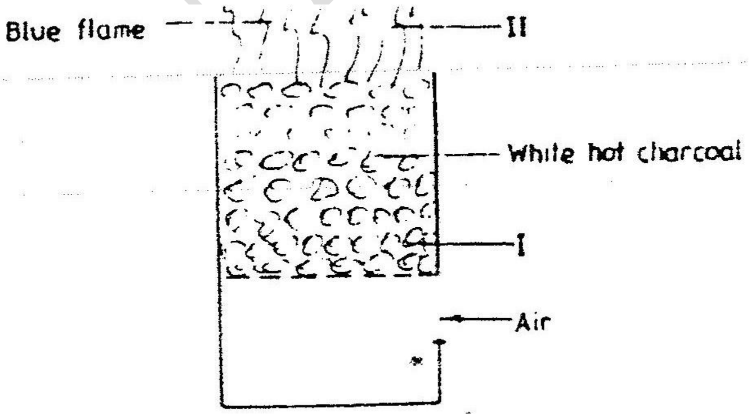
$Fe_3O_4(s) + 4CO(g) \rightarrow 3Fe(s) + 4CO_2(g)$
(brown when hot/cool (grey)

These reaction are used during the extraction of many metals from their ore.

9. Carbon (II) oxide is a pollutant. Explain.

Carbon(II)oxide is highly poisonous/toxic.It preferentially combine with haemoglobin to form stable carboxyhaemoglobin in the blood instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood causing nausea , coma then death.

10.The diagram below show a burning charcoal stove/burner/jiko. Use it to answer the questions that follow.



Explain the changes that take place in the burner

Explanation

Charcoal stove has air holes through which air enters. Air oxidizes carbon to carbon(IV)oxide gas at region I. This reaction is exothermic($-\Delta H$) producing more heat.

Chemical equation

$C(s) + O_2(g) \rightarrow CO_2(g)$

Carbon(IV)oxide gas formed rises up to meet more charcoal which reduces it to Carbon(II)oxide gas.

Chemical equation

$2CO_2 (g) + O_2(g) \rightarrow 2CO (g)$

At the top of burner in region II, Carbon (II)oxide gas is further oxidized to Carbon(IV)oxide gas if there is plenty of air but escape if the air is limited poisoning the living things around.

Chemical equation

$2CO (g) + O_2(g) \rightarrow 2CO_2 (g)$
(excess air)

11.Describe the test for the presence of carbon(II)oxide gas.

Experiment

Burn/Ignite the pure sample of the gas. Pass/Bubble the products into lime water/Calcium hydroxide .

Observation

Colourless gas burns with a blue flame. A white precipitate is formed that dissolve on further bubbling of the products.

Chemical equation

$2CO (g) + O_2(g) \rightarrow 2CO_2 (g)$ (gas burns with blue flame)

Chemical equation

$Ca(OH)_2 (aq) + CO_2 (g) \rightarrow CaCO_3 (s) + H_2O(l)$

Chemical equation

$CO_2 (g) + CaCO_3 (s) + H_2O(l) \rightarrow Ca(HCO_3)_2 (aq)$

12. State the main uses of carbon (II)oxide gas.

(i) As a fuel /water gas

(ii)As a reducing agent in the blast furnace for extracting iron from iron ore(Magnetite/Haematite)

(iii)As a reducing agent in extraction of Zinc from Zinc ore/Zinc blende

(iv) As a reducing agent in extraction of Lead from Lead ore/Galena

(v) As a reducing agent in extraction of Copper from Copper iron sulphide/Copper pyrites.

(iii)Carbonate(IV) (CO_3^{2-})and hydrogen carbonate(IV(HCO_3^-)

1.Carbonate (IV) (CO_3^{2-}) are normal salts derived from carbonic(IV)acid (H_2CO_3) and hydrogen carbonate (IV) (HCO_3^-) are acid salts derived from carbonic(IV)acid.

Carbonic(IV)acid(H_2CO_3) is formed when carbon(IV)oxide gas is bubbled in water. It is a dibasic acid with two ionizable hydrogens.

$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$

$H_2CO_3(aq) \rightarrow H^+(aq) + HCO_3^- (aq)$

2.Carbonate (IV) (CO_3^{2-}) are insoluble in water **except** Na_2CO_3 , K_2CO_3 and $(NH_4)_2CO_3$

3.Hydrogen carbonate (IV) (HCO_3^-) are soluble in water. Only five hydrogen carbonates exist. **$NaHCO_3$, $KHCO_3$,NH_4HCO_3 $Ca(HCO_3)_2$ and $Mg(HCO_3)_2$**

$Ca(HCO_3)_2$ and $Mg(HCO_3)_2$ exist only in aqueous solutions.

3.The following experiments show the effect of heat on Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) salts:

Experiment

In a clean dry test tube place separately about 1.0 of the following:

Zinc(II)carbonate(IV), sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) ammonium carbonate(IV), potassium hydrogen carbonate(IV), Lead(II)carbonate(IV), Iron(II)carbonate(IV), and copper(II)carbonate(IV). Heat each portion gently the strongly. Test any gases produced with lime water.

Observation

(i)Colorless droplets form on the cooler parts of test tube in case of sodium carbonate(IV) and Potassium carbonate(IV).

(ii)White residue/solid left in case of sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) and potassium hydrogen carbonate(IV).

(iii)Colour changes from blue/green to black in case of copper(II)carbonate(IV).

(iv) Colour changes from green to brown/yellow in case of Iron (II)carbonate(IV).

(v) Colour changes from white when cool to yellow when hot in case of Zinc (II) carbonate(IV).

(vi) Colour changes from yellow when cool to brown when hot in case of Lead (II) carbonate(IV).

(vii)Colourless gas produced that forms a white precipitate with lime water in all cases.

Explanation

1. Sodium carbonate(IV) and Potassium carbonate(IV) exist as hydrated salts with 10 molecules of water of crystallization that condenses and collects on cooler parts of test tube as a colourless liquid.

Chemical equation

$Na_2CO_3 .10H_2O(s) \;\; -> \;\; Na_2CO_3 (s) \;\; + \;\; 10H_2O(l)$
$K_2CO_3 .10H_2O(s) \;\; -> \;\; K_2CO_3 (s) \;\; + \;\; 10H_2O(l)$

2. Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) salts decompose on heating except Sodium carbonate(IV) and Potassium carbonate(IV).

(a) Sodium hydrogen carbonate(IV) and Potassium hydrogen carbonate(IV) decompose on heating to form sodium carbonate(IV) and Potassium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

$2NaHCO_3 (s) \;\; -> \;\; Na_2CO_3 (s) \;\; + \;\; H_2O(l) + CO_2 (g)$
(white) (white)
$2KHCO_3 (s) \;\; -> \;\; K_2CO_3 (s) \;\; + \;\; H_2O(l) + CO_2 (g)$
(white) (white)

(b) Calcium hydrogen carbonate(IV) and Magnesium hydrogen carbonate(IV) decompose on heating to form insoluble Calcium carbonate(IV) and Magnesium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

$Ca(HCO_3)_2 (aq) \;\; -> \;\; CaCO_3 (s) \;\; + \;\; H_2O(l) + CO_2 (g)$
(Colourless solution) (white)
$Mg(HCO_3)_2 (aq) \;\; -> \;\; MgCO_3 (s) \;\; + \;\; H_2O(l) + CO_2 (g)$
(Colourless solution) (white)

(c) Ammonium hydrogen carbonate(IV) decompose on heating to form ammonium carbonate(IV) .Water and carbon(IV)oxide gas are also produced.

Chemical equation

$2NH_4HCO_3$ (s) -> $(NH_4)_2CO_3$ (s) + H_2O(l) + CO_2 (g)
(white) (white)

(d)All other carbonates decompose on heating to form the metal oxide and produce carbon(IV)oxide gas e.g.

Chemical equation

$MgCO_3$ (s) -> MgO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$BaCO_3$ (s) -> BaO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$CaCO_3$ (s) -> CaO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$CuCO_3$ (s) -> CuO (s) + CO_2 (g)
(blue/green solid) (black solid)

Chemical equation

$ZnCO_3$ (s) -> ZnO (s) + CO_2 (g)
(white solid) (white solid when cool/
Yellow solid when hot)

Chemical equation

$PbCO_3$ (s) -> PbO (s) + CO_2 (g)
(white solid) (yellow solid when cool/
brown solid when hot)

4.The following experiments show the presence of Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^{-}) ions in sample of a salt:

(a)Using Lead(II) nitrate(V)

I. Using a portion of salt solution in a test tube .add four drops of Lead(II)nitrate(V)solution.Preserve.

Observation	inference
White precipitate/ppt	CO_3^{2-} ,SO_3^{2-} ,SO_4^{2-} ,Cl^{-} -

II. To the preserved solution ,add six drops of dilutte nitric(V)acid. Preserve.

Observation	inference
White precipitate/ppt persists	SO_4^{2-} ,Cl^{-}
White precipitate/ppt dissolves	CO_3^{2-} ,SO_3^{2-}

II. To the preserved sample(that forms a precipitate),heat to boil.

Observation	inference

White precipitate/ppt persists	SO_4^{2-}
White precipitate/ppt dissolves	Cl^-

II. To the preserved sample(that do not form a precipitate),add three drops of acidified potassium manganate(VII)/lime water

Observation	inference
Effervescence/bubbles/fizzing colourless gas produced Acidified $KMnO_4$ decolorized/no white precipitate on lime water	SO_3^{2-}
Effervescence/bubbles/fizzing colourless gas produced Acidified $KMnO_4$ not decolorized/ white precipitate on lime water	CO_3^{2-}

Experiments/Observations:

(b)Using Barium(II)nitrate(V)/ Barium(II)chloride

I. To about 5cm3 of a salt solution in a test tube add four drops of Barium(II) nitrate (V) / Barium(II)chloride. Preserve.

Observation	Inference
White precipitate/ppt	SO_4^{2-} , SO_3^{2-} , CO_3^{2-} ions

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

Observation	Inference
White precipitate/ppt persists	SO_4^{2-} , ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves	SO_3^{2-} , CO_3^{2-} , ions

III.To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

Observation	Inference
(i)acidified potassium manganate(VII)decolorized (ii)Orange colour of acidified potassium dichromate(VI) turns to green	SO_3^{2-} ions

Observation 2

Observation	Inference
(i)acidified potassium manganate(VII) not decolorized (ii)Orange colour of acidified potassium dichromate(VI) does not turns to green	CO_3^{2-} ions

Explanations

Using Lead(II)nitrate(V)

(i)Lead(II)nitrate(V) solution reacts with chlorides(Cl^-), Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Lead(II)chloride, Lead(II)sulphate(VI), Lead(II) sulphate (IV) and Lead(II)carbonate(IV).

Chemical/ionic equation:

$Pb^{2+}(aq) + Cl^-(aq) \rightarrow PbCl_2(s)$

$Pb^{2+}(aq) + SO_4^{2+}(aq) \rightarrow PbSO_4(s)$

$Pb^{2+}(aq) + SO_3^{2+}(aq) \rightarrow PbSO_3(s)$

$Pb^{2+}(aq) + CO_3^{2+}(aq) \rightarrow PbCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Lead(II)chloride and Lead(II)sulphate(VI) do not react with the acid and thus their white precipitates remain/ persists.

- Lead(II) sulphate (IV) and Lead(II)carbonate(IV) reacts with the acid to form **soluble** Lead(II) nitrate (V) and produce/effervesces/fizzes/bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$PbSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + SO_2(g)$

$PbCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + CO_2(g)$

(iii)When Lead(II)chloride and Lead(II)sulphate(VI) are heated/warmed;

- Lead(II)chloride dissolves in hot water/on boiling(recrystallizes on cooling)

- Lead(II)sulphate(VI) do not dissolve in hot water thus its white precipitate persists/remains on heating/boiling.

(iv)When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green. **Carbon(IV)oxide will not.**

Chemical equation:

$5SO_3^{2-}(aq) + \underset{\text{(purple)}}{2MnO_4^-(aq)} + 6H+(aq) \rightarrow 5SO_4^{2-}(aq) + \underset{\text{(colourless)}}{2Mn^{2+}(aq)} + 3H_2O(l)$

$3SO_3^{2-}(aq) + \underset{\text{(Orange)}}{Cr_2O_7^{2-}(aq)} + 8H+(aq) \rightarrow 3SO_4^{2-}(aq) + \underset{\text{(green)}}{2Cr^{3+}(aq)} + 4H_2O(l)$

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

Using Barium(II)nitrate(V)/ Barium(II)Chloride

(i)Barium(II)nitrate(V) and/ or Barium(II)chlo**ride** solution reacts with Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Barium(II)sulphate(VI), Barium(II) sulphate (IV) and Barium(II)carbonate(IV).

Chemical/ionic equation:

$Ba^{2+}(aq) + SO_4^{2+}(aq) \rightarrow BaSO_4(s)$

$Ba^{2+}(aq) + SO_3^{2+}(aq) \rightarrow BaSO_3(s)$

$Ba^{2+}(aq) + CO_3^{2+}(aq) \rightarrow BaCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Barium (II)sulphate(VI) do not react with the acid and thus its white precipitates remain/ persists.

- Barium(II) sulphate (IV) and Barium(II)carbonate(IV) reacts with the acid to form **soluble** Barium(II) nitrate (V) and produce /effervesces /fizzes/ bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$BaSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + SO_2(g)$

$BaCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + CO_2(g)$

(iii) When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green. **Carbon(IV)oxide will not.**

Chemical equation:

$5SO_3^{2-}(aq) + \underset{\text{(purple)}}{2MnO_4^-(aq)} + 6H+(aq) \rightarrow 5SO_4^{2-}(aq) + \underset{\text{(colourless)}}{2Mn^{2+}(aq)} + 3H_2O(l)$

$3SO_3^{2-}(aq) + \underset{\text{(Orange)}}{Cr_2O_7^{2-}(aq)} + 8H+(aq) \rightarrow 3SO_4^{2-}(aq) + \underset{\text{(green)}}{2Cr^{3+}(aq)} + 4H_2O(l)$

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

(iii) Sodium carbonate(IV) (Na_2CO_3)

(a)Extraction of sodium carbonate from soda ash

Sodium carbonate naturally occurs in Lake Magadi in Kenya as Trona.trona is the double salt ; sodium sesquicarbonate. $NaHCO_3 .Na_2CO_3 .H_2O$.It is formed from the volcanic activity that takes place in Lake Naivasha, Nakuru ,Bogoria and Elementeita .All these lakes drain into Lake Magadi through underground rivers. Lake Magadi has no outlet.

Solubility of Trona decrease with increase in temperature.High temperature during the day causes trona to naturally crystallize .It is mechanically scooped/dredged/dug and put in a furnace.

Inside the furnace, trona decompose into soda ash/sodium carbonate.

Chemical equation

$\underset{(trona)}{2NaHCO_3 .Na_2CO_3 .H_2O(s)} \rightarrow \underset{(soda\ ash)}{3Na_2CO_3(s)} + 5H_2O(l) + CO_2(g)$

Soda ash is then bagged and sold as Magadi soda.It is mainly used:

(i)in making glass to lower the melting point of raw materials (sand/SiO_2 from 1650°C and CaO from 2500°C to around 1500°C)

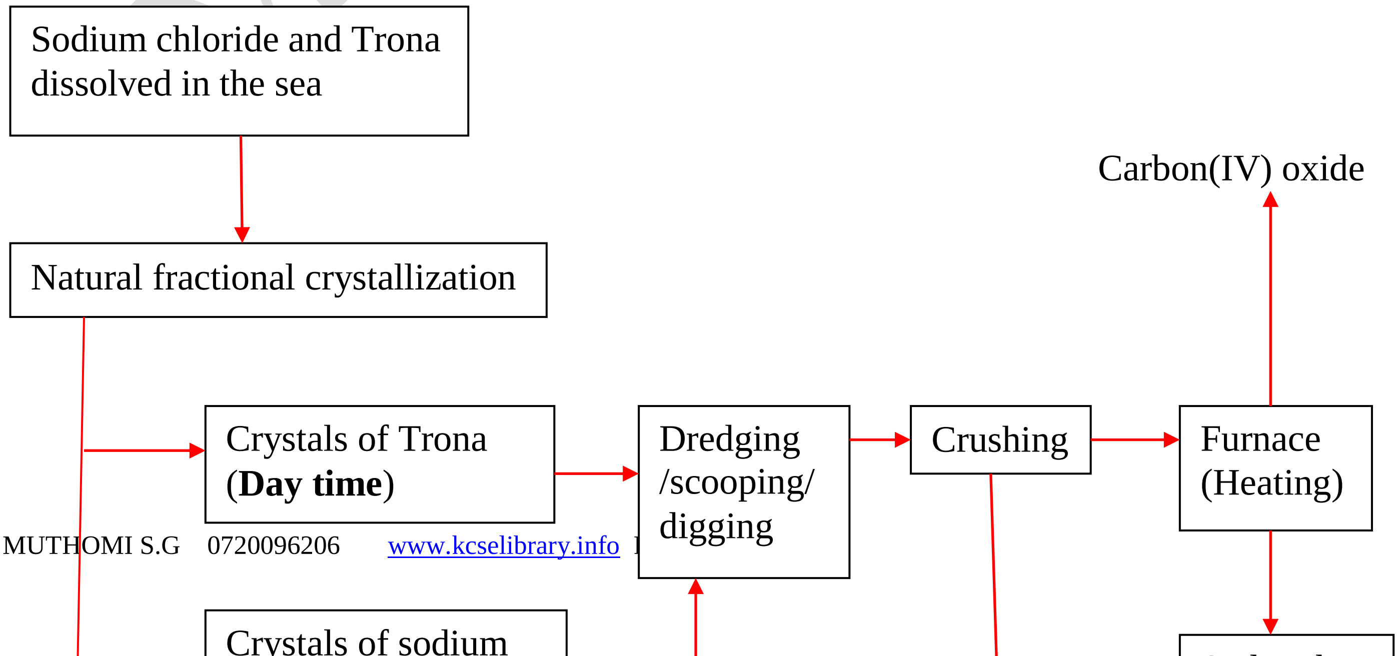
(ii)in softening hard water

(iii)in the manufacture of soapless detergents.

(iv)Swimming pool "pH increaser"

Sodium chloride is also found dissolved in the lake. Solubility of sodium chloride decrease with decreases in temperature/ sodium chloride has lower solubility at lower temperatures. When temperatures decrease at night it crystallize out .The crystals are then mechanically dug/dredged /scooped then packed for sale as animal/cattle feeds and seasoning food.

Summary flow diagram showing the extraction of Soda ash from Trona



b)The Solvay process for industrial manufacture of sodium carbonate(IV)

(i)Raw materials.

-**Brine** /Concentrated Sodium chloride from salty seas/lakes.
-**Ammonia** gas from Haber.
-**Limestone** /Calcium carbonate from chalk /limestone rich rocks.
-**Water** from rivers/lakes.

(ii)Chemical processes

Ammonia gas is passed **up** to meet a **downward** flow of sodium chloride solution / brine to form **ammoniated** brine/**ammoniacal** brine **mixture** in the **ammoniated brine chamber**

The ammoniated brine mixture is then pumped up, atop the carbonator/ solvay tower.

In the carbonator/ solvay tower, ammoniated brine/ammoniacal brine mixture slowly trickle down to meet an upward flow of carbon(IV)oxide gas.

The carbonator is shelved /packed with quartz/broken glass to

(i) reduce the rate of flow of ammoniated brine/ammoniacal brine mixture.

(ii)increase surface area of the liquid mixture to ensure a lot of ammoniated brine/ammoniacal brine mixture react with carbon(IV)oxide gas.

Insoluble sodium hydrogen carbonate and soluble ammonium chloride are formed from the reaction.

Chemical equation

$CO_2(g) + H_2O(l) + NaCl\ (aq) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

The products are then filtered. **Insoluble** sodium hydrogen carbonate forms the **residue** while soluble ammonium chloride forms the **filtrate**.

Sodium hydrogen carbonate itself can be used:

(i) as baking powder and preservation of some soft drinks.
(ii) as a buffer agent and antacid in animal feeds to improve fibre digestion.
(iii) making dry chemical fire extinguishers.

In the Solvay process Sodium hydrogen carbonate is then heated to form Sodium carbonate/soda ash, water and carbon (IV) oxide gas.

Chemical equation

$2NaHCO_3\ (s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

Sodium carbonate is stored ready for use in:

(i) during making glass/lowering the melting point of mixture of sand/SiO_2 from 1650°C and CaO from 2500°C to around 1500°C
(ii) in softening hard water
(iii) in the manufacture of soapless detergents.
(iv) swimming pool “pH increaser”.

Water and carbon(IV)oxide gas are recycled back to the ammoniated brine/ammoniacal brine chamber.

More carbon(IV)oxide is produced in the kiln/furnace. Limestone is heated to decompose into Calcium oxide and carbon(IV)oxide.

Chemical equation

$CaCO_3 (s) \rightarrow CaO(s) + CO_2(g)$

Carbon(IV)oxide is recycled to the carbonator/solvay tower. Carbon (IV)oxide is added water in the **slaker** to form Calcium hydroxide. This process is called **slaking**.

Chemical equation

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

Calcium hydroxide is mixed with ammonium chloride from the carbonator/solvay tower in the **ammonia regeneration chamber** to form Calcium chloride , water and more ammonia gas.

Chemical equation

$Ca(OH)_2 (aq) + 2NH_4Cl (aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

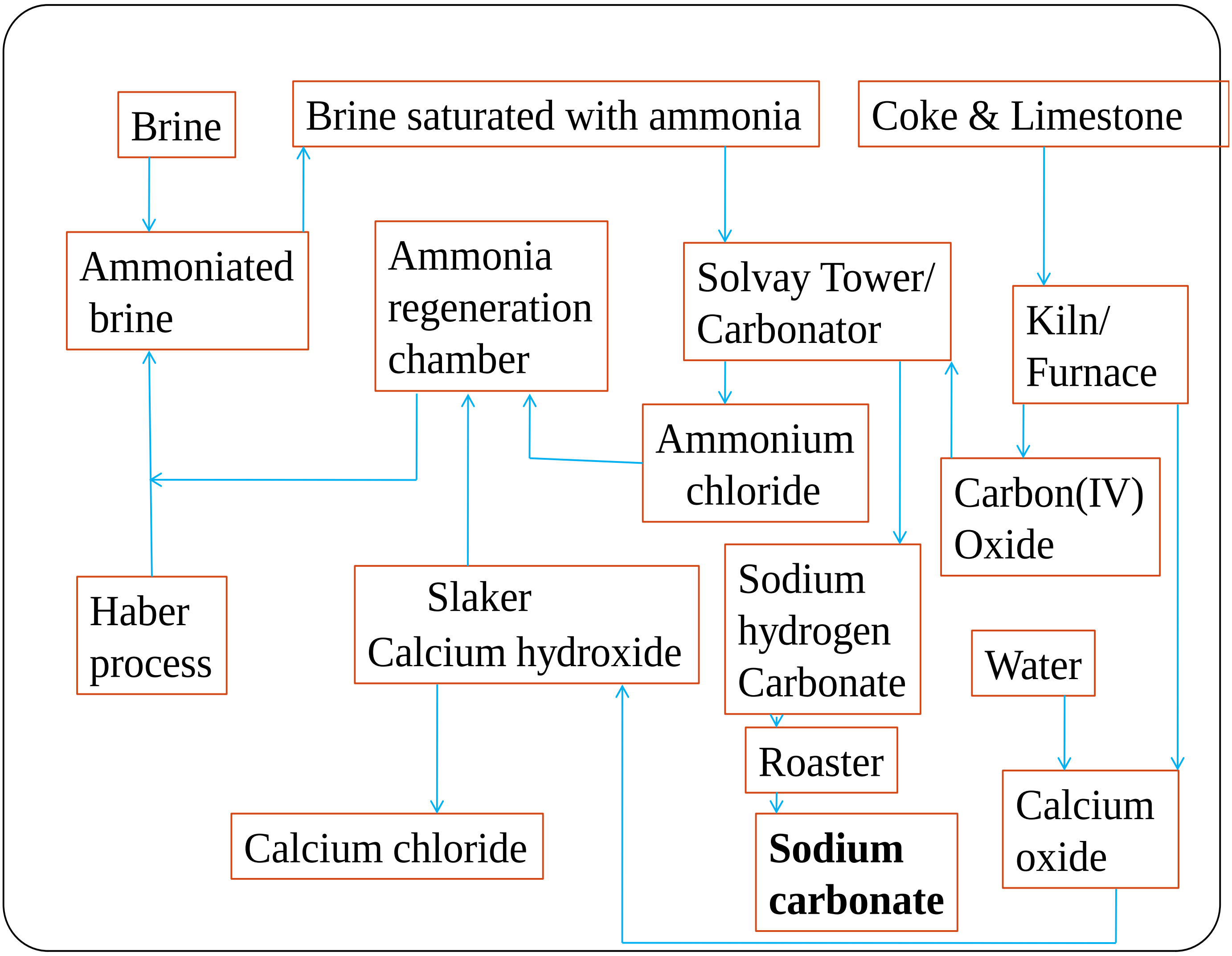
$NH_3(g)$ and $H_2O(l)$ are recycled.

Calcium chloride may be used:

(i)as drying agent in the school laboratory during gas preparation (except ammonia gas)

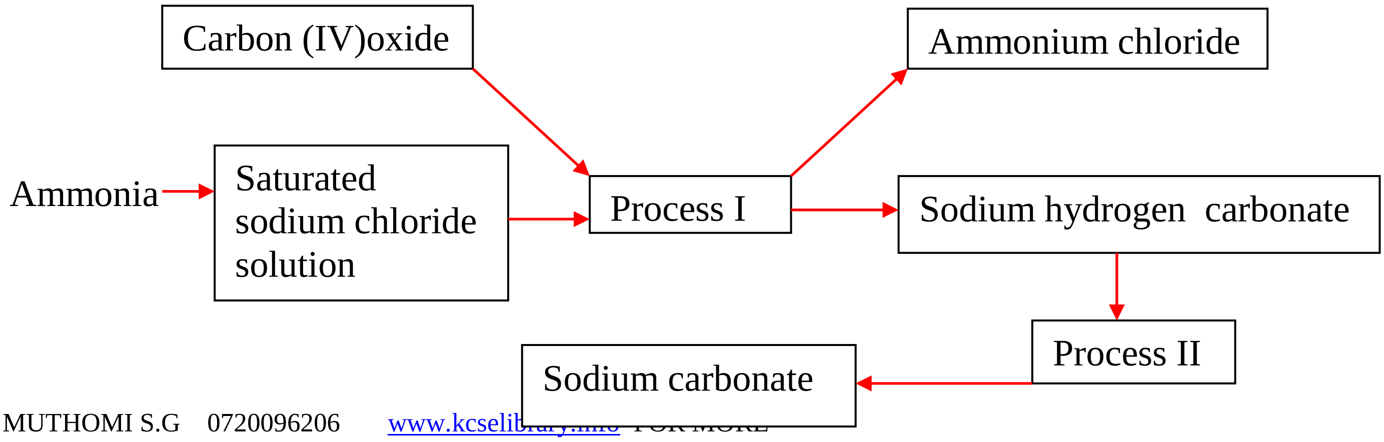
(ii)to lower the melting point of solid sodium chloride / rock salt salts during the Downs process for industrial extraction of sodium metal.

Detailed Summary flow diagram of Solvay Process



Practice

1. The diagram below shows part of the Solvay process used in manufacturing sodium carbonate. Use it to answer the questions that follow.



(a)Explain how Sodium Chloride required for this process is obtained from the sea.
Sea water is pumped /scooped into shallow pods. Evaporation of most of the water takes place leaving a very concentrated solution.

(b)(i) Name process:

I. Filtration

II. Decomposition

(ii) Write the equation for the reaction in process:

Process I

Chemical equation

$CO_2(g) + H_2O(l) + NaCl\ (aq) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

Process II

Chemical equation

$2NaHCO_3\ (s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

(c)(i) Name two substances recycled in the solvay process

Ammonia gas , Carbon(IV)Oxide and Water.

(ii)Which is the by-product of this process?

Calcium(II)Chloride /$CaCl_2$

(iii)State two uses that the by-product can be used for:

1. As a drying agent in the school laboratory preparation of gases.
2. In the Downs cell/process for extraction of Sodium to lower the melting point of rock salt.

(iv)Write the chemical equation for the formation of the by-products in the Solvay process.

Chemical equation

$Ca(OH)_2\ (aq) + 2NH_4Cl\ (aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

(d)In an experiment to determine the % purity of Sodium carbonate produced in the Solvay process ,2.15g of the sample reacted with exactly 40.0cm3 of 0.5M Sulphuric(VI)acid.

(i)Calculate the number of moles of sodium carbonate that reacted.

Chemical equation

$Na_2CO_3\ (aq) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + CO_2(g) + H_2O(l)$

Mole ratio $Na_2CO_3 : H_2SO_4 \Rightarrow 1:1$

$$\text{Moles } H_2SO_4 = \frac{\text{Molarity x Volume}}{1000} \Rightarrow \frac{0.5 \text{ x } 40.0}{1000} = \underline{\mathbf{0.02\ Moles}}$$

Moles of Na_2CO_3 = **<u>0.02 Moles</u>**

(ii)Determine the % of sodium carbonate in the sample.

Molar mass of Na_2CO_3 = **106g**

Mass of Na_2CO_3 = moles x Molar mass => 0.02 x 106 = **<u>2.12 g</u>**

% of Na_2CO_3 = $\frac{(2.12\text{ g} \times 100)}{2.15}$ = **<u>98.6047%</u>**

(e) State two uses of soda ash.

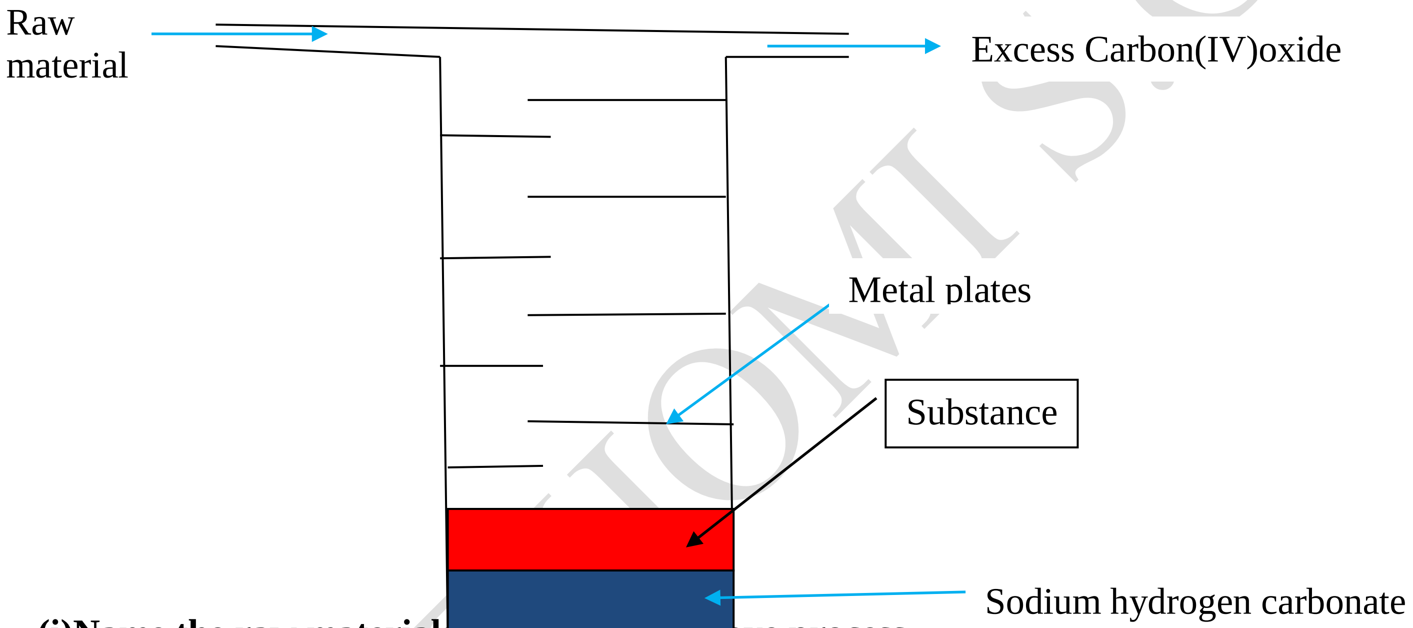
(i) during making glass/lowering the melting point of mixture of sand/SiO_2 from 1650°C and CaO from 2500°C to around 1500°C

(ii) in softening hard water

(iii) in the manufacture of soapless detergents.

(iv) swimming pool "pH increaser".

(f)The diagram below shows a simple ammonia soda tower used in manufacturing sodium carbonate .Use it to answer the questions that follow:



(i)Name the raw materials needed in the above process

- Ammonia
- Water
- Carbon(IV)oxide
- Limestone
- Brine/ Concentrated sodium chloride

(ii)Identify substance A

Ammonium chloride /NH_4Cl

(iii) Write the equation for the reaction taking place in:

I.Tower.

<u>Chemical equation</u>

$CO_2(g) + NaCl\ (aq) + H_2O(l) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

II. Production of excess carbon (IV)oxide.

<u>Chemical equation</u>

$CaCO_3\ (s) \rightarrow CaO(s) + CO_2(g)$

III. The regeneration of ammonia

Chemical equation

$Ca(OH)_2$ (aq) +$2NH_4Cl$ (aq) -> $CaCl_2$(s) + $2NH_3$(g) + H_2O(l)

(iv)Give a reason for having the circular metal plates in the tower.

-To slow the downward flow of brine.

-To increase the rate of dissolving of ammonia.

-To increase the surface area for dissolution

(v)Name the gases recycled in the process illustrated above.

Ammonia gas , Carbon(IV)Oxide and Water.

2. Describe how you would differentiate between carbon (IV)oxide and carbon(II)oxide using chemical method.

Method I

-Bubble both gases in lime water/$Ca(OH)_2$

-white precipitate is formed if the gas is carbon (IV) oxide

- **No** white precipitate is formed if the gas is carbon (II) oxide

Method II

-ignite both gases

- Carbon (IV) oxide does **not** burn/ignite

- Carbon (II) oxide burn with a blue non-sooty flame.

Method III

-Lower a burning splint into a gas containing each gas separately.

-burning splint is extinguished if the gas is carbon (IV) oxide

-burning splint is **not** extinguished if the gas is carbon (II) oxide.

3.Using Magnesium sulphate(VI)solution ,describe how you can differentiate between a solution of sodium carbonate from a solution of sodium hydrogen carbonate

-Add Magnesium sulphate(VI) solution to separate portions of a solution of sodium carbonate and sodium hydrogen carbonate in separate test tubes

-White precipitate is formed in test tube containing sodium carbonate

-No white precipitate is formed in test tube containing sodium hydrogen carbonate.

Chemical equation

Na_2CO_3 (aq) +$MgSO_4$ (aq) -> Na_2SO_4 (aq) + $MgCO_3$(s)
(white ppt)

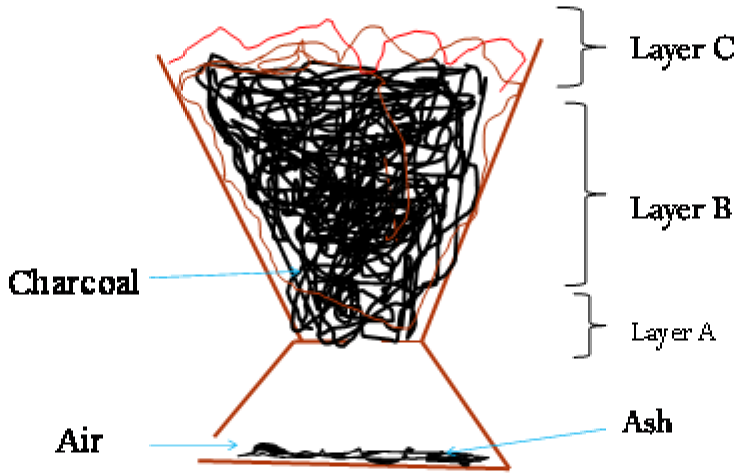
Ionic equation

CO_3^{2-} (aq) + Mg^{2+} (aq) -> $MgCO_3$(s)
(white ppt)

Chemical equation

$2NaHCO_3$ (aq) +$MgSO_4$ (aq) -> Na_2SO_4 (aq) + $Mg(HCO_3)_2$ (aq)
(colourless solution)

4. The diagram below shows a common charcoal burner .Assume the burning take place in a room with sufficient supply of air.



(a)Explain what happens around:

(i)Layer A

Sufficient/excess air /oxygen enter through the air holes into the burner .It reacts with/oxidizes Carbon to carbon(IV)oxide

Chemical equation

$C(s) + O_2(g) \rightarrow CO_2(g)$

(ii)Layer B

Hot carbon(IV)oxide rises up and is reduced by more carbon/charcoal to carbon (II)oxide.

Chemical equation

$C(s) + CO_2(g) \rightarrow 2CO(g)$

(ii)Layer C

Hot carbon(II)oxide rises up and burns with a blue flame to be oxidized by the excess air to form carbon(IV)oxide.

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

(b)State and explain what would happen if the burner is put in an enclosed room.

The hot poisonous /toxic carbon(II)oxide rising up will not be oxidized to Carbon(IV)oxide.

(c)Using a chemical test , describe how you would differentiate two unlabelled black solids suspected to be charcoal and copper(II)oxide.

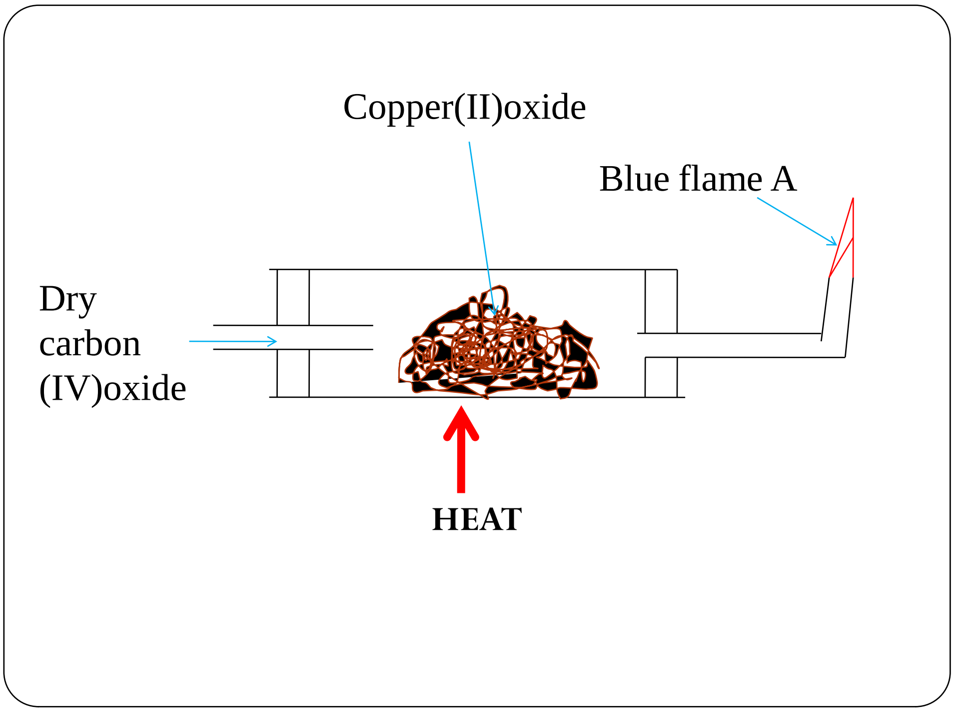
Method I

-Burn/Ignite the two substances separately.
-Charcoal burns with a blue flame
- Copper(II)oxide does not burn

Method II

-Add dilute sulphuric(VI)acid/Nitric(V)acid/Hydrochloric acid separately.
-Charcoal does not dissolve.
- Copper(II)oxide dissolves to form a colourless solution.

5. Excess Carbon(II)oxide was passed over heated copper(II)oxide as in the set up shown below for five minutes.



(a)State and explain the observations made in the combustion tube.

Observation

Colour change from black to brown

Explanation

Carbon (II)oxide reduces black copper(II)oxide to brown copper metal itself oxidized to Carbon(IV)oxide.

Chemical equation

$CO(g)$ + $CuO(s)$ -> $Cu(s)$ + $CO_2(g)$
(black) (brown)

(b) (i)Name the gas producing flame A

Carbon(II)oxide

(ii)Why should the gas be burnt?

It is toxic/poisonous

(iii)Write the chemical equation for the production of flame A

$2CO(g)$ + $O_2(g)$ -> $2CO_2(g)$

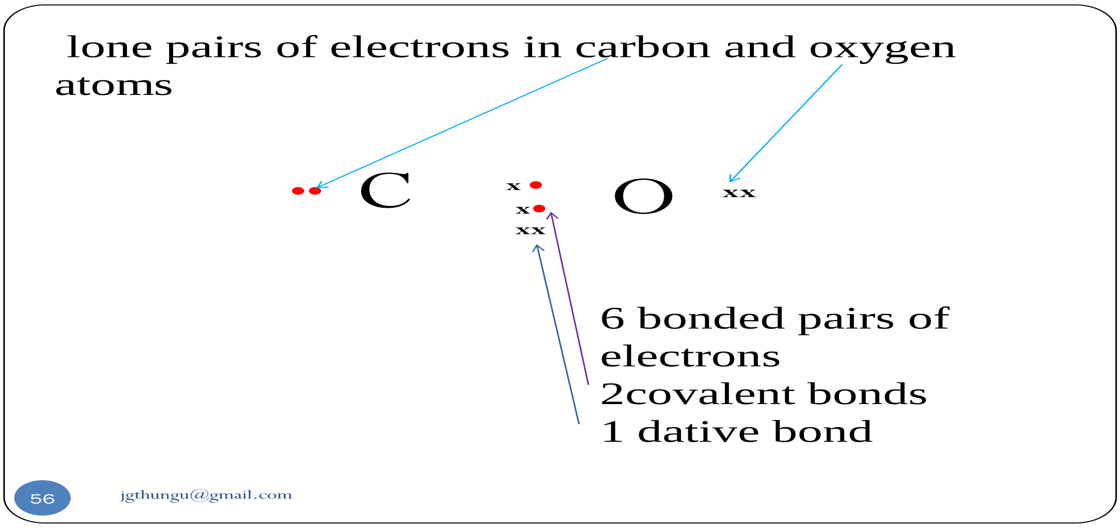
(c)State and explain what happens when carbon(IV)oxide is prepared using Barium carbonate and dilute sulphuric(VI)acid.

Reaction starts then stops after sometime producing small/little quantity of carbon(IV)oxide gas.

Barium carbonate react with dilute sulphuric(VI)acid to form insoluble Barium sulphate(VI) that cover/coat unreacted Barium carbonate stopping further reaction to produce more Carbon(IV)oxide.

(d) Using dot () and cross(x) to represent electrons show the bonding in a molecule of :

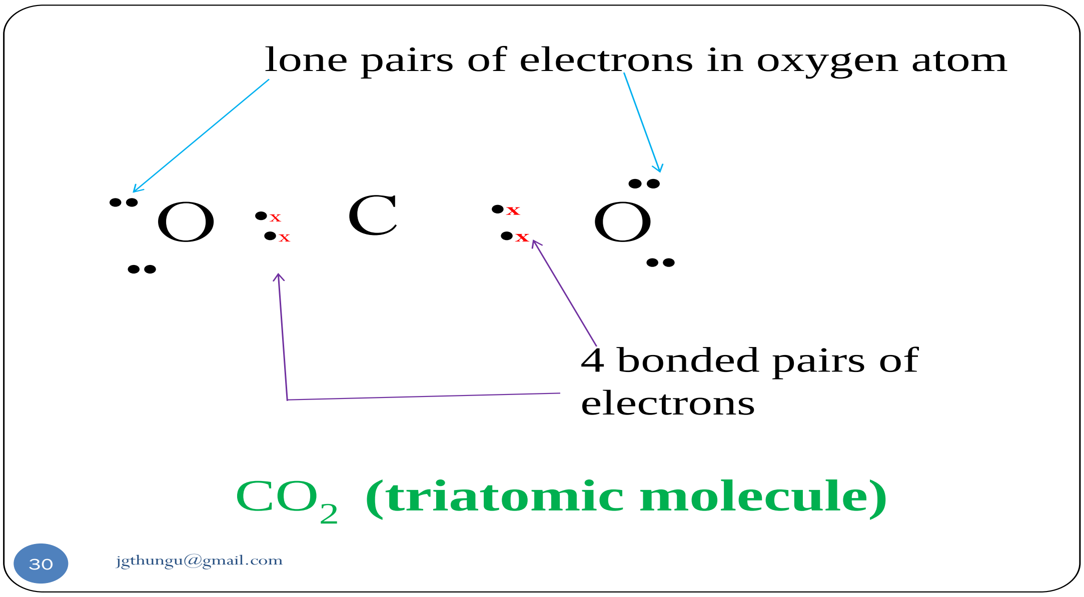
(i) Carbon(II)oxide



lone pairs of electrons in carbon and oxygen atoms

6 bonded pairs of electrons
2covalent bonds
1 dative bond

(ii) Carbon(IV)Oxide.



lone pairs of electrons in oxygen atom

4 bonded pairs of electrons

CO_2 (triatomic molecule)

(e) Carbon (IV)oxide is an environmental pollutant of global concern. Explain.

-It is a green house gas thus causes global warming.

-It dissolves in water to form acidic carbonic acid which causes “acid rain”

(f)Explain using chemical equation why lime water is used to test for the presence of Carbon (IV) oxide instead of sodium hydroxide.

Using lime water/calcium hydroxide:

- a visible white precipitate of calcium carbonate is formed that dissolves on bubbling excess Carbon (IV) oxide gas

<u>Chemical equation</u>

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(white precipitate)

$CaCO_3(aq) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$

Using sodium hydroxide:

- **No** precipitate of sodium carbonate is formed Both sodium carbonate and sodium hydrogen carbonate are soluble salts/dissolves.

Chemical equation

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

(No white precipitate)

$Na_2CO_3 (s) + H_2O(l) + CO_2 (g) \rightarrow 2NaHCO_3 (s)$

(g)Ethan-1,2-dioic acid and methanoic acid may be used to prepare small amount of carbon(II)oxide in a school laboratory.

(i) Explain the modification in the set up when using one over the other.

Before carbon(II)oxide is collected:

-when using methanoic acid, **no** concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

-when using ethan-1,2-dioic acid, concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

(ii)Write the equation for the reaction for the formation of carbon(II)oxide from:

I.Methanoic acid.

Chemical equation $HCOOH(aq) \rightarrow CO(g) + H_2O(l)$

II. Ethan-1,2-dioic acid

Chemical equation $HOOCCOOH(aq) \rightarrow CO_2(g)+CO(g)+H_2O(l)$

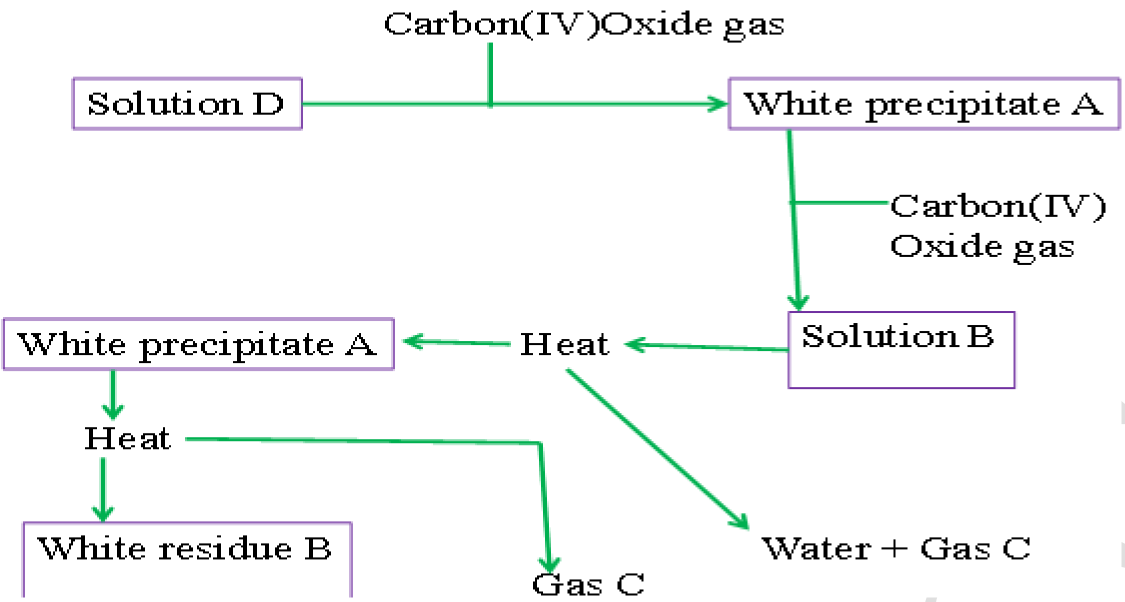
(h)Both carbon(II)oxide and carbon(IV)oxide affect the environment. Explain why carbon(II)oxide is more toxic/poisonous.

-Both gases are colourless,denser than water and odourless.

-Carbon(II)oxide is preferentially absorbed by human/mammalian haemoglobin when inhaled forming stable carboxyhaemoglobin instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood leading to suffocation and quick death. --Carbon(IV)oxide is a green house gas that increases global warming.

-Carbon(II)oxide is readily oxidized to carbon(IV)oxide

6.Study the flow chart below and use it to answer the questions that follow.



(a)Name:

(i)the white precipitate A

Calcium carbonate

(ii) solution B

Calcium hydrogen carbonate

(iii) gas C

Carbon(IV)oxide

(iv) white residue B

Calcium oxide

(v) solution D

Calcium hydroxide/lime water

(b)Write a balanced chemical equation for the reaction for the formation of:

(i) the white precipitate A from solution D

Chemical equation

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(ii) the white precipitate A from solution B

Chemical equation

$Ca(HCO_3)_2(aq) \rightarrow CO_2(g) + CaCO_3(s) + H_2O(l)$

(iii) solution B from the white precipitate A

Chemical equation

$CO_2(g) + CaCO_3(s) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$

(iv) white residue B from the white precipitate A

Chemical equation

$CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

(iv) reaction of white residue B with water

Chemical equation

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