Preface
The structure of the atom is extensively covered here. The periodic table is the arrangement of atoms of elements based on their atomic structure. Emphasis on the trends across and down the periodic table of atoms is important for the teacher facilitator. This work should be covered before "**Chemical bonding and structure"**. The two complements each other in understanding **periodicity**. Candidate-user preparing for any secondary level chemistry from members of

A.<u>ATOMIC STRUCTURE</u>

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

(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

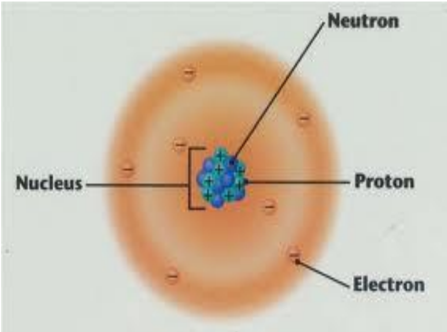
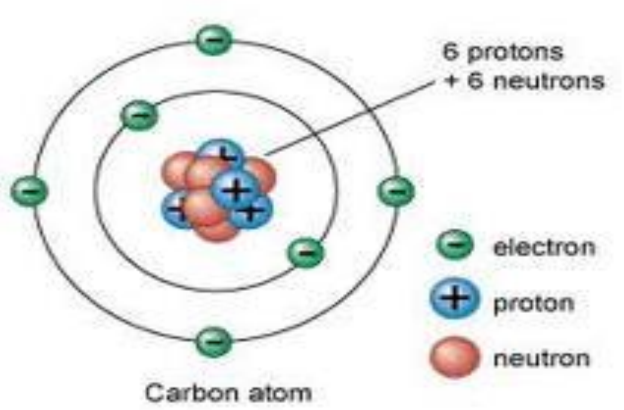


Diagram showing the relative positions of protons, electrons and neutrons in an atom of Carbon



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
atomic number $\quad {}^{m}_{n}X \quad$ symbol of element

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	${}^{39}_{19}K$ ${}^{40}_{19}K$	19 19	19 19	20 21	19 19	39 40

	$^{41}_{19}K$	19	19	22	19	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% $^{35}_{17}Cl$ and 25% $^{37}_{17}Cl$ isotopes. Calculate the relative atomic mass of Chlorine.

Working

100 atoms of chlorine contains 75 atoms of $^{35}_{17}Cl$ isotopes

100 atoms of chlorine contains 75 atoms of $^{37}_{17}Cl$ isotopes

Therefore;

RAM of chlorine = (75/100 x 35) + 25/100 x 37 = **35.5**

Note that:

Relative atomic mass has no units

More atoms of chlorine exist as $^{35}_{17}Cl$(75%) than as $^{37}_{17}Cl$(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% $^{39}_{19}K$, 0.01% $^{40}_{19}K$, 6.89% $^{41}_{19}K$,

Working
100 atoms of potassium contains 93.1 atoms of $^{39}_{19}K$ isotopes
100 atoms of potassium contains 0.01 atoms of $^{40}_{19}K$ isotopes
100 atoms of potassium contains 6.89 atoms of $^{41}_{19}K$ 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 no units
More atoms of potassium exist as $^{39}_{19}K$ (93.1%) therefore RAM is nearer to the more abundant $^{39}_{19}K$ isotope.

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

Working
100 atoms of Neon contains 90.92 atoms of $^{20}_{10}Ne$ isotopes
100 atoms of Neon contains 0.26 atoms of $^{21}_{10}Ne$ isotopes
100 atoms of Neon contains 8.82 atoms of $^{22}_{10}Ne$ 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 no units
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 higher the relative atomic mass, the heavier 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

Practice examples drawing electronic configurations

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**

Commented [s1]:

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 1^{st} energy levels
- **8** electron in the 2^{nd} energy levels
- **5** electron in the 3^{rd} 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 1^{st} energy levels
- **8** electron in the 2^{nd} energy levels
- **6** electron in the 3^{rd} 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 1^{st} energy levels
- **8** electron in the 2^{nd} energy levels
- **7** electron in the 3^{rd} 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 1^{st} energy levels
- **8** electron in the 2^{nd} energy levels
- **8** electron in the 3^{rd} 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**

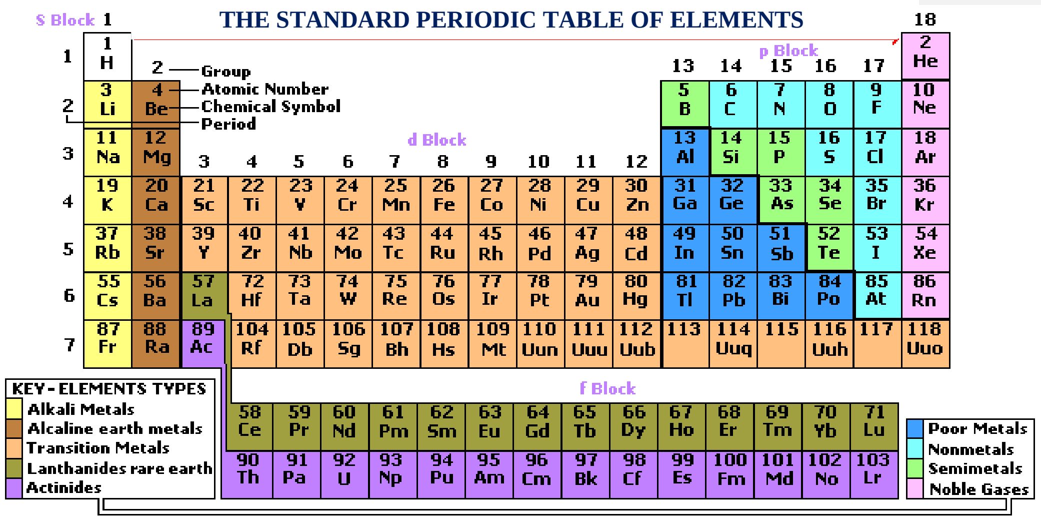
d) $^{1}_{1}H$
Electron structure **1**: => 1 electron in outer energy level thus **Group I**

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.



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	2:8:3		2:8		
structure	(unstable)		(stable)		

3. $^{23}_{11}$Na

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

4. $^{24}_{12}$Mg

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

5. $^{16}_{8}$O

	O	+	2e	->	O^{2-}
	(atom)		(2 electrons gained/acquired)		(anion)
Electron	2:6				2:8
structure	(unstable)				(stable)

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

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

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

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

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

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

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

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

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

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

Practice examples

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: **Al_2O_3**

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: **Na_2O**

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**: <u>CaO</u>**
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**: <u>PbO_2</u>**
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**: <u>PbO</u>**
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**: <u>Fe_2O_3</u>**
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**: $\underline{FeSO_4}$**
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**: $\underline{CuSO_4}$**
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**: $\underline{Al_2(SO_4)_3}$**
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**: $\underline{Al\ (NO_3)_3}$**
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.

<u>Practice examples</u>

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 <u>equal</u> to product side

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

One atom of Cl in HCl on the reactant side is <u>not equal</u> 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

$\mathbf{H_2O_2 \rightarrow H_2O + O_2}$

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.

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

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

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

(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

$\mathbf{KMnO_4 + HCl \rightarrow KCl + MnCl_2 + H_2O + Cl_2}$

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.

$NaOH(aq)$ + $HCl(aq)$ -> $NaCl\ (aq)$ + $H_2O(l)$

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

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

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

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

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

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

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

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

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

$Mg(s)$ + $\underline{2}H_2O(g)$ -> $MgO(s)$ + $H_2(g)$

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

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

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

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

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

2**$C_2H_2(g)$ + $\underline{5}O_2(g)$ -> $\underline{4}CO_2(g)$ + $\underline{2}H_2O(l)$**

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).

<u>Atomic and ionic radius</u>
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.

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

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

$$Na_2O(s) \quad + \quad H_2O(l) \quad -> \quad 2NaOH(aq)$$

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

$$2NaOH(aq) \quad + \quad CO_2(g) \quad -> \quad Na_2CO_3(g) \quad + \quad H_2O(l)$$

(ii)Burning in air/oxygen

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

$$4Li\ (s) \quad + \quad O_2(g) \ -> \quad 2Li_2O(s)$$

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

$$4Na\ (s) \quad + \quad O_2(g) \ -> \quad 2Na_2O(s)$$

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

$$2Na\ (s) \quad + \quad O_2(g) \ -> \quad Na_2O_2\ (s)$$

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

$$4K\ (s)+ \quad O_2(g) \ -> \quad 2K_2O\ (s)$$

(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	Very vigorous

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

Explanation

Alkali metals are less dense than water. They therefore float in water.They react with water to form a strongly alkaline solution of their hydroxides and producing hydrogen gas. The rate of this reaction increase down the group. i.e. Potassium is more reactive than sodium .Sodium is more reactive than Lithium. The reactivity increases as electropositivity increases of the alkali increases. This is because as the atomic radius increases , the ease of donating/losing outer electron increase during chemical reactions.

Chemical equations

$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$
$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
$2Rb(s) + 2H_2O(l) \rightarrow 2RbOH(aq) + H_2(g)$
$2Cs(s) + 2H_2O(l) \rightarrow 2CsOH(aq) + H_2(g)$
$2Fr(s) + 2H_2O(l) \rightarrow 2FrOH(aq) + H_2(g)$

Reactivity increase down the group

(iv) Reaction with chlorine:

Experiment

Cut about 0.5g of sodium into a deflagrating spoon with a lid cover. Introduce it on a Bunsen flame until it catches fire. Quickly and carefully lower it into a gas jar containing dry chlorine to cover the gas jar.
Repeat with about 0.5g of Lithium.
Caution: This experiment should be done in fume chamber because chlorine is poisonous /toxic.

Observation

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

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

Chemical equations

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

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

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

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

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

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

Reactivity increase down the group

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

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

Some **uses** of alkali metals include:
(i)Sodium is used in making sodium cyanide for extracting gold from gold ore.
(ii)Sodium chloride is used in seasoning food.
(iii)Molten mixture of sodium and potassium is used as coolant in nuclear reactors.
(iv)Sodium is used in making sodium hydroxide used in making soapy and soapless detergents.

(v)Sodium is used as a reducing agent for the extraction of titanium from Titanium(IV)chloride.
(vi)Lithium is used in making special high strength glasses
(vii)Lithium compounds are used to make dry cells in mobile phones and computer laptops.

Group II elements: Alkaline earth metals

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

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

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

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

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

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

Element	Symbol	Atomic number	Atomic radius(nM)	Ionic radius(nM)
Beryllium	Be	4	0.089	0.031
Magnesium	Mg	12	0.136	0.065

Calcium	Ca	20	0.174	0.099

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

Electropositivity

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

Ionization energy

For alkaline earth metals the 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

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

<u>Experiment</u>

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.

<u>Observations</u>

-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.

<u>Explanations</u>

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.

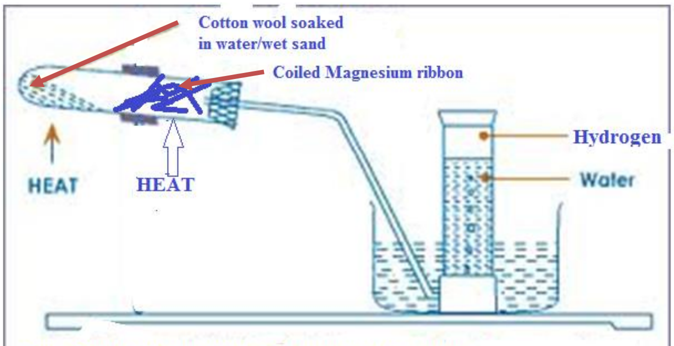
$Ca(s) + 2H_2O (l) \rightarrow Ca(OH)_2(aq) + H_2 (g)$

(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 in tetrachloromethane	-101	-35
Bromine	Br_2	Poor	Insoluble/soluble in tetrachloromethane	7	59
Iodine	I_2	Poor	Insoluble/soluble in tetrachloromethane	114	sublimes

Chemical properties

(i)Displacement

Experiment

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

Add 5 drops of chlorine water to each test tube:

Repeat with 5 drops of bromine water instead of chlorine water

Observation

Using Chlorine water

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

-Coloured Solution formed.

Using Bromine water

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

-Coloured Solution formed.

Explanation

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

Chlorine is more electronegative than bromine and iodine.

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

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

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

Table showing the displacement of the halogens

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

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

Chemical /ionic equations

With Fluorine

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

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

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

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

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

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

With chlorine

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

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

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

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

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**	**SiL_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)