TERMS COMMONLY USED IN INORGANIC CHEMISTRY

- Atomic radius
- Ionic radius
- Charge density
- Ionization energy
- Electron affinity
- Electronegativity
- Shielding / screening effect
- Nuclear charge
- Complex ion
- Hydrogen bond
- Lattice energy
- Coordinate bond
- Polarization
- Polarizing power

Definitions:

1. **Screening effect / shielding effect:**

Is the tendency of the inner electrons on inner energy levels to prevent the attraction of outer elections by the nucleus. Screening effect increases with increase in the number of energy levels; therefore screening effect increases down a given group; whereas screening effect remains almost constant across a given period because the number of energy levels remains constant.

2. **Nuclear charge:**

This is the ability of the nucleus to attract electrons towards itself. Nuclear charge depends on the number of protons in the nucleus. Nuclear charge increases down a group and across a given period because the number of protons increase.

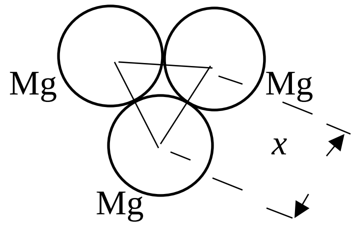
Note:

The difference between nuclear charge and screening effect is known as effective nuclear charge. Effective nuclear charge increases across a given period because the increase in nuclear charge outweighs the increase in screening effect.

Effective nuclear charge / attraction decreases down a given group because the increase in screening effect outweighs the increase in the nuclear charge.

3. **Atomic radius:**

For metals, it is referred to as metallic radius metallic radius is half the inter nuclear distance between two nuclei of adjacent atoms in a lattice for example;



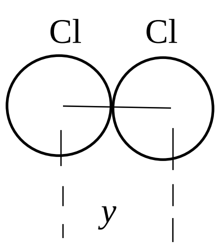
Atomic radius of Mg $=\frac{x}{2}$.

For non-metals, it is referred to as covalent radius.

Is half the inter nuclear distance between two atoms covalently bonded to each other by single covalent bond.

For example;

Chlorine



Atomic radius of chlorine $=\frac{y}{2}$.

Note:

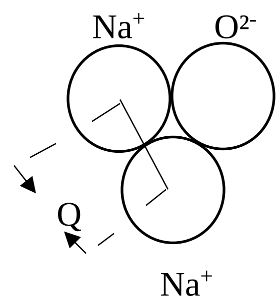
(i) Generally, atomic radius is half the distance between two nuclei of two atoms which are bonded together.

(ii) The atomic radius of an uncombined atom cannot be defined strictly because of the uncertain boundary of electron clouds. But the distance between the nuclei of chemically combined atoms can be measured accurately.

(iii) Ionic radius is half the distance between the nuclei of two adjacent ions in an ionic crystal.

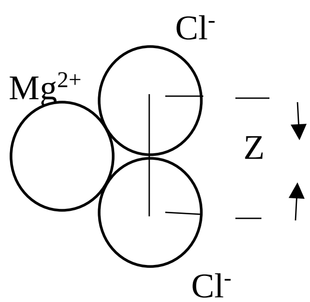
For example;

Ionic radius of sodium ion (Na^+) can be obtained from sodium oxide (Na_2O) as follows;



Ionic radius of $Na^+ = \frac{Q}{2}$.

Ionic radius of chloride ion (Cl) can be obtained from the structure of mgcl2 as follows.



Ionic radius of $Cl^- = \frac{Z}{2}$

(iv) All cations are smaller than the neutral atoms from which they are formed. For example;

Sodium ion is smaller than sodium atom. This is because in the formation of cations, electrons in the external shells of atoms are removed, this decreases the screening effect, while the nuclear charge remains the same because protons do not change and so the effective nuclear charge increases and therefore, the cation become smaller.

(v) All anions are larger than the corresponding neutral atoms.

For example;

The chloride ion is larger than chloride atom. This is because electrons have been added to complete the external shell, this increases the screening effect, the nuclear charge remains the same because the number of protons are unchanged; the effective nuclear charge decreased making the anion to become bigger.

Variation of atomic radius down a given group.

For example;

Group (II): Elements

Alkaline earth metals

Element	**Atomic radius**
Be	0.112
Mg	0.160
Ca	0.197
Sr	0.215
Ba	0.219

Group (VII): Elements

Halogens

Element	**Atomic radius**
Fluorine	0.72
Chlorine	0.99
Bromine	1.14
Iodine	1.33

Trend: Atomic radius increases down the group.

Explanation:

Nuclear charge increases as protons are being added to the nucleus as you move from one element to the next down the group, screening effect increases because of extra shell of electrons is being added. The increase in screening effect outweighs the increase in nuclear charge making the effective nuclear charge to reduce hence increase in atomic radius.

Variation of atomic radius across period 3.

Element in period 3	Na	Mg	Al	Si	P	S	Cl
Atomic radius (nm)	0.157	0.136	0.125	0.117	0.11	0.104	0.099

Atomic radius reduces / decreases across period 3.

Explanation:

Nuclear charge increases because the number of protons in the nucleus increases across a period, screening effect remains almost constant because the number of energy levels remain the same. The increase nuclear charge outweighs the screening effect, effective nuclear attraction for elections increase, making the atomic radius to decrease.

4. **Charge density:**

This is the ratio of **ionic charge** to **ionic radius.**

$$\text{Charge density} = \frac{\text{Ionic charge}}{\text{Ionic radius}}$$

Example:

Calculate the charge density of the sodium ion given that its ionic radius is 0.785nm.
Solution:

$$\text{Charge density of } Na^{+} = \frac{+1}{0.785}$$

$$= +\ 1.274 nm^{-1}$$

Exercise:

Calculate the charge density of the following ions given the ionic radius in brackets.

(i) Aluminum ion (0.05nm)
(ii) Magnesium ion (0.065nm)
(iii) Lithium ion (0.06)
(iv) Beryllium ion (0.051nm)

5. **Polarizing power:**

Polarizing power of an ion is a measure of its ability to attract or pull electrons from another atom or ion with negative charge.
Note:

An ion with a **high charge density** has **high polarizing** power.

Example:

Aluminum ion (Al^{3+}) has a higher charge density and higher polarizing power than sodium ion.

6. Electronegativity:

This is the power of an atom in a molecule to attract bonding electrons to itself in a covalent bond.

$$\text{Electronegativity} = \frac{\text{Effective nuclear charge}}{\text{Covalent atomic radius}}$$

That is to say:

Electronegativity is directly proportional to effective nuclear charge and inversely proportional to atomic radius.
If there is a sufficient difference in electronegativity between two atoms, polarization will be virtually complete, and ionization will have resulted.

For example:

A : B A and B have the same electronegativity.

$A^{\delta-} : B^{\delta+}$ A more electronegative than B

$A^{-} : B^{+}$ A much more electronegative than B.

Some electronegativity values of some elements are given below:

Halogens	F	Cl	Br	I
Electronegativity	4.0	3.0	2.8	2.5

Alkali metals	Li	Na	K	Rb	Cs
Electronegativity value	1.0	0.9	0.8	0.8	0.7

Period 3 Element	Na	Mg	Al	Si	P	S	Cl
Electronegativity value	0.9	1.25	1.5	1.8	2.15	2.5	2.8

The most electronegative atoms are:

Fluorine with electronegative value 4.0.

Oxygen with electronegative value 3.5.

Nitrogen with electronegative value 3.0.

Variation of electronegativity down a group.

Trend:

Electronegativity decreases down a group.

Explanation:

Atomic radius increases down a group, screening effect increases because an extra shell of electrons is added, nuclear charge increases because protons are being added, effective nuclear charge reduces making electronegativity to decrease because it's directly proportional to effective nuclear charge and inversely proportional to atomic radius.

Variation of electronegativity do across a period.

Trend:

Electronegativity increases across a period.

Explanation:

Atomic radius reduces across a period, screening effect remains almost the same because there is no change in the number of energy levels, nuclear charge increases because extra proton is added to the nucleus, effective nuclear charge increases, making electronegativity to increase.

7. **Electropositivity:**

This is the tendency of an atom to donate electrons to form positively charged ions.

Electro positivity increases down a group because atomic radius increases, screening effect increases because extra shell of electrons is added, nuclear charge increases because protons added to the nucleus, the increase in screening effect outweighs the increase in nuclear charge, effective nuclear attraction for outermost electrons reduce, making them to be easily lost.

Trend:

Electro positivity decreases across a period.

Explanation:

Note:

- The most electropositive elements are metals.
- The most electronegative elements are non-metals.
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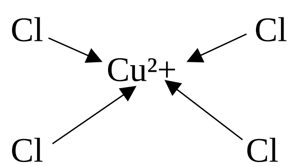
8. **Complex ion:**

Is an ion in which atoms or group of atoms with lone pairs of electrons associate with a central atom which is usually a metal.

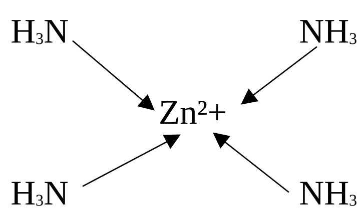
Examples of complex ions;

$Cu(NH_3)_4^{2+}$ structure

$CuCl_4^{2-}$



$Zn(NH_3)_4^{2+}$



9. **Covalent bond:**

This is a bond formed between two atoms with a small difference in electronegativity by sharing of electrons between the atoms.
For example;
H-Br, Cl – Cl, C – H, N – H

10. **Ionic bond:**

This is a bond formed between highly electropositive atom and a highly electronegative atom by transfer of electrons.
For example;

$Na^+ Cl^-$

11. Hydrogen bond:

This is a bond formed by the dipole-diploe attraction between a hydrogen atom of one molecule and a highly electronegative atom of another molecule.

$H^{\delta+} - F^{\delta-}$ .... $H^{\delta+} - F^{\delta-}$

Hydrogen bond

$H^{\delta+} - F^{\delta-}$ .... $H^{\delta+} - O^{\delta-} - H$

Hydrogen bond

12. Ionization energy (I.E):

This can be defined as the minimum amount of energy required to remove an electron from a gaseous atom to form a unipositively gaseous ion.

$M(g) \longrightarrow M^+(g) + e \quad \Delta H$ = First ionization energy of element M.

Where M represents a symbol of an element in the periodic table, ΔH is the amount energy.

OR

It can be defined as the minimum amount of energy required to remove one mole of electrons from on mole of gaseous atoms to form one mole of unipositively gaseous ions.

Ionization energy is the same as first ionization energy.

After removal of the first one mole of electrons, the energy required to remove the second mole of electrons is known as second ionization energy.

$M^+(g) \longrightarrow M^{2+}(g) + e$

Therefore, an element can have first, second, third, fourth …and n^{th} ionization energies until all the electrons have been removed.

The first ionization energy < second ionization energy < 3^{rd} I.E < 4^{th} I.E < 5^{th} I.E etc.

This is because as electrons are removed, the proton / electron ratio increases, this increases the nuclear charge, the effective nuclear attraction for the remaining electrons increases. This result into increase in ionization energy.

Quiz:

Explain / Why the second ionization energy is greater than the first ionization energy.

If electrons of an atom are removed from the same energy level, the difference between successive ionization energies is not big, however, there is always a largest energy difference if two electrons are being removed from different energy levels.

Illustration:

Sodium

1^{st} I.E = 494 kδmol-1

2^{nd} I.E = 4500 kδmol-1

} Very large difference because the electrons are being removed from difference energy levels.

Magnesium

1^{st} I.E = 736 kδmol-1

2^{nd} I.E = 1450 kδmol-1

3^{rd} I.E = 1450 kδmol-1

Aluminum

1^{st} I.E = 577 kδmol-1

2^{nd} I.E = 1820 kδmol-1

3^{rd} I.E = 2740 kδmol-1

4^{th} I.E = 11600 kδmol-1

Question 1:

The table below shows the ionization energies in kδmol-1 for elements A, B, C, D and E.

Element	**1^{st} I.E**	**2^{nd} I.E**	**3^{rd} I.E**	**4^{th} I.E**
A	500	4600	6900	9500
B	740	1500	7700	10500
C	630	1600	3000	48000
D	900	1800	14800	21000
E	580	1800	2700	11600

a) Which one of these elements is an alkali metal? Give a reason for your answer.
b) State two elements which belong to the same group in the periodic table.
c) Select an element which is most likely to form a caution with +3?
d) Write the formula of the oxide of E.

Question 2:

The first six successive ionization energies of the elements of a given period of the periodic table are shown in the table.

Element	**1st I.E**	**2nd I.E**	**3rd I.E**	**4th I.E**	**5th I.E**	**6th I.E**
A	1000	2260	3340	4540	6990	8490
B	786	1580	3230	4360	16000	20000
C	1520	2660	3950	5770	7240	8790
D	577	1820	2740	11600	14800	18400
E	1060	1900	2920	4960	6280	21200
F	1260	2300	3850	5150	6540	9330
G	494	4560	6940	9540	13400	16600
H	736	1450	7740	10500	13600	18000

a) State and explain the trend in the successive ionization energies.
b) State with a reason which element represents.
   (i) An alkaline earth metal.
   (ii) Noble gas
c) Write equation to show how
   (i) The third electron is removed from A.
   (ii) The second electron is removed from H.
d) Write the formula of;
   (i) Oxide of D
   (ii) Chloride of E
   (iii) Nitride of B

Note:

(i) Metals generally have low first ionization energy.
(ii) Non-metals generally have high first ionization energy.
(iii) Noble gases (or inert gases) have very high first ionization energy.

Metals are normally classified as electropositive elements because they can easily don not electrons and form positive ions.
Metals are also reducing agents because they can easily donate electrons due to the low ionization energy.

$Na\ (g) \longrightarrow Na^{+}\ (g) + e$

$Mg\ (g) \longrightarrow Mg^{+}\ (g) + e$

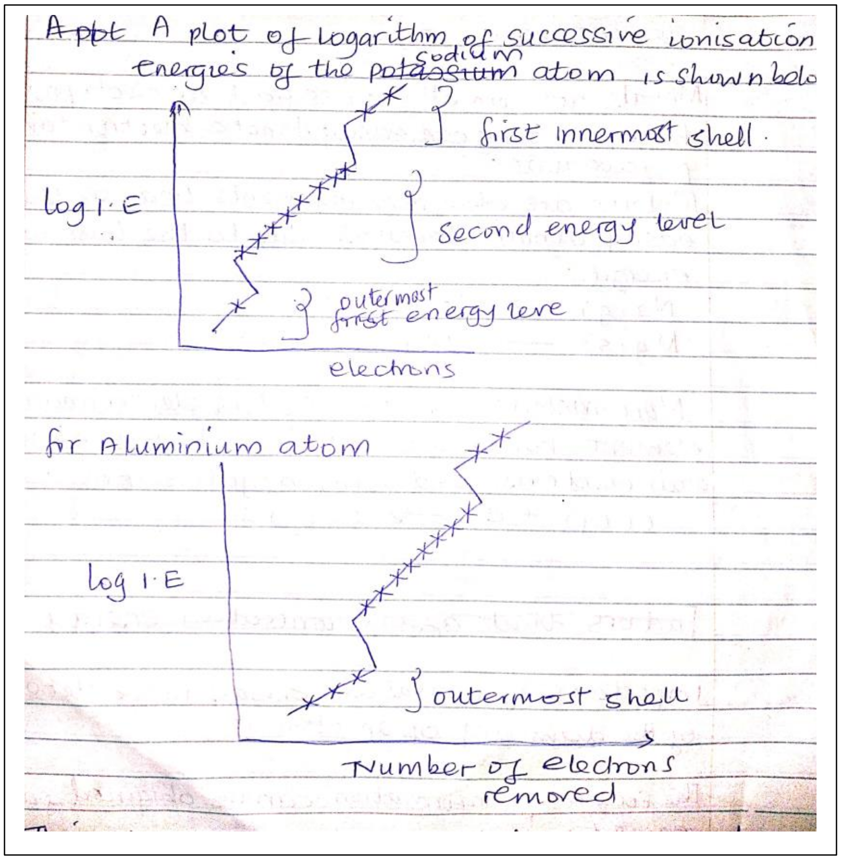
Non – metals are classified as electronegative elements because they have a high tendency to gain electrons and form negative ions.

$Cl\ (g) + e \longrightarrow Cl\ (g)$

Importance of ionization energy in the determination of the chemistry of an element.

The following information can be obtained from ionization energy.

- Atomic number of an element.
  The successive ionization energies indicate the number of electrons present in an atom.
  For example;
  Aluminum has 13 successive ionization energies indicating that the atomic number of Aluminum is 13.
- Determination of metallic character of an element metals have low first ionization energy (less than 800).
- Arrangement of electrons and distribution of energy levels.



This shows that aluminum belongs to period 3 because it has three energy levels, shown by the three distinct breaks, and has a valence of 3 because it has three electrons in the outer most energy level.

Quiz:

Explain why there is a very large energy difference between the 3rd I.E and 4th I.E of aluminum.

Quiz:

The first eight ionization energies of an elements B are shown below;

	1st	2nd	3rd	4th	5th	6th	7th	8th
Ionization energies	786	1580	3230	4360	16000	20000	23600	29100

a) To which group in the periodic table does B belong? Give a reason for your answer.
b) Plot a graph of logarithm of ionization energy against number of electrons.
c) Explain why the first ionization of B is smaller than its second ionization energy.

Factors that affect ionization energy:

1. Nuclear charge. The higher the nuclear charge, the higher the ionization energy.
2. Screening effect. The higher the screening effect, the lower the ionization energy.
3. Atomic radius.
4. Electronic configuration. When an electron is being removed from a sub shell which is either full of electrons or half filled, the amount of energy required to remove an electron will be high.

Quiz:

Explain why the ionization energy of magnesium is higher than that of Aluminum.

Variation of first ionization energy across.

Period 3:

Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
I.E (kδmol-1)	500	737	577	786	1012	999	1255	1521
Electronic Configuration								

Trend:

The first ionization energy generally increases across period 3 from sodium to argon.

Explanation:

The nuclear charge increases because a proton is being added to the nucleus, screening effect increases slightly (remains almost constant) because an electron is being added on the same energy level. The increase in nuclear charge outweighs the screening effect, effective nuclear attraction for the electrons increases resulting into an increase in energy required to remove an electron.

However, the first ionization energy of aluminum is lower than that of magnesium because of their electronic configuration.

Mg = $1s^2 2s^2 2p^6 3s^2 3p^1$

For magnesium, 3s outer subshell has two electrons, it is completely filled and thermodynamically stable, requires a lot of energy to remove the first electron.

For Aluminum, 3p outer subshell has only one electron, it is neither half-filled nor completely filled, and therefore, unstable, requires less energy to remove an electron.

Also, the first ionization energy of phosphorus is higher than that of sulphur because of their electronic configuration.

P = $1s^2 2s^2 2p63s^2 3p^3$, S = $1s^2 2s^2 sp_6 3s^2 3p^4$.

For phosphorus, it has 3p outermost subshell having three electrons, it is half filled and thermodynamically stable, requires much more energy to remove the first electron from phosphorous.

For sulphur, it has 3p outermost subshell having four electrons, it is neither half-filled nor completely filled, making it thermodynamically unstable, requires much less energy to remove the first electron from sulphur.

Question:

The table below shows the first ionization energy of period 2 elements.

Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Atomic number	3	4	5	6	7	8	9	10
First ionization energy ($kjmol^{-1}$)	520	899	801	1086	1403	1310	1681	2080

(a) Plot a graph of ionization energy against atomic number. (4 marks)

(b) Explain the shape of the graph. (10 marks)

Variation of ionization energy down the group.

For example;

Group II:

Element	**First ionization energy**
Be	899
Mg	737
Ca	590
Sr	543
Ba	504

Trend:

The first ionization energy decreases down the group.

Explanation:

The nuclear charge increases due to extra protons added, screening effect increases because of extra shell of electrons added, the increase in screening effect outweighs the increase in nuclear charge, and the effective nuclear attraction for the electrons reduces making ionization energy to reduce.

13. Electron affinity:

This is the heat given out when an electron is added to a gaseous atom to form a uninegatively gaseous ion.

$Y\ (g) + e \longrightarrow Y^-(g)$

$O\ (g) + e \longrightarrow O^-\ (g)$

$Cl\ (g) + e \longrightarrow Cl^-\ (g)$

Or

It is the heat given out when one mole of electrons is added to one mole of gaseous atoms to form one mole of negatively charged gaseous ions.

Electron affinity is an exothermic process.

Note:

The first electron affinity of a given gaseous atom is always an exothermic process because there is no repulsion between the existing electrons and an electron added.

For example:

$S\ (g) + e \longrightarrow S^-\ (g)$ -217 $kjmol^{-1}$

$O\ (g) + e \longrightarrow O^-\ (g)$ -300 $kjmol^{-1}$

The second electron affinity of a given gaseous atom always an endothermic process / positive because there is repulsion between the gaseous ion and the second electron added. Heat has to be supplied for the second electron to be accepted making it positive.

$O^-\ (g) + e \longrightarrow O^{2-}\ (g)$ +20 $kjmol^{-1}$

$S^-\ (g) + e \longrightarrow S^{2-}\ (g)$ +150 $kjmol^-$

Factors that affect electron affinity.

- Nuclear charge.

  Electron affinity increases with increase in nuclear charge and decreases with decrease in nuclear charge.

- Screening effect.
  Electron affinity decreases with increase in screening effect and it increases with decrease in screening effect.
- Atomic radius.
  When the atomic radius is small, the electron affinity is very high.
- Electron configuration of atom.
  If an electron is being added to a half filled or completely filled subshell, less energy will be given off.

Variation of electron affinity across period 3.

Element	Na	Mg	Al	Si	P	S	Cl
Atomic number	11	12	13	14	15	16	17
Electron affinity	-80	+67	-30	-135	-60	-200	-364

Trend:

The first electron affinity generally increases across the period.

Explanation:

This is due to decrease in atomic radius, the nuclear charge increases because a proton is added to the nucleus, screening effect remains almost constant because an electron is being added on the same energy level, the increase in nuclear attraction outweighs the screening effect, the effective nuclear attraction increases leading to increase electron affinity.

The first electron affinity of magnesium is positive (endothermic).

Mg = $1s^22s^22p^63s^2$

The 3s in magnesium has two electrons; it is completely filled and stable, the electron added experiences more repulsion than nuclear attraction.

The electron affinity of phosphorous is lower than the expected because of its electronic configuration.

P = $1s^22s^22p^63s^23p^3$.

The 3p in phosphorous has three electrons, it is half filled and stable, the electron added experiences more repulsion than nuclear attraction.

Variation of electron affinity down the group.

For example;

Group (VII) elements.

Element	Electron affinity
Fluorine	-323
Chlorine	-364
Bromine	-324
Iodine	-295

The electron affinity generally decreases from chlorine to iodine, due to increase in atomic radius. This is because the screening effect increases because an extra shell of electrons is being added, nuclear charge increases because protons are being added to the nucleus, the increase in screening effect outweighs the increase in nuclear charge, effective nuclear attraction reduces hence a decrease in amount of energy given out.

The electron affinity of fluorine is unexpectedly low because fluorine atom has the smallest atomic radius, electron added experience more repulsion from existing lone pair of electrons than nuclear attraction.

Note

The chemistry of the first element in each group is different from the rest of the group members.

For example;

1. The chemistry of lithium is different from the rest of the group(I) elements because;
   (i) Lithium atom has the smallest atomic radius.
   (ii) Lithium atom has the highest electronegativity value.

Some of the properties in which lithium differs the rest of the group members include,

- Lithium hydrogen carbonate exists in solution state.
- It burns in air to form ionic nitride.
- It burns in oxygen form a normal oxide.

- Lithium carbonate is decomposed by heat.
- Lithium hydroxide is also decomposed by heat.
- Lithium nitrate is decomposed by heat to form metal oxide, nitrogen dioxide and oxygen gas.

2. The chemistry of beryllium is different from the rest of the group (II) elements because,

- Beryllium atom has the smallest atomic radius.
- Beryllium has the highest electronegativity value amongst the group members.
- Beryllium lacks the d-orbital.

3. The chemistry of carbon is different from the rest of group (iv) members because;

- Carbon has the smallest atomic radius.
- Carbon has the highest electronegativity value.
- Carbon lacks the 2 d orbitals.

Some of the properties in which carbon is different from the rest of group(IV) members include;

- Carbon forms gaseous oxides.
- Carbon forms multiple bonds with itself and with other non-metals.
- Carbon has ability to catenate (bond to itself to form long chains).
- Carbon has a maximum valence of 4.

4. The chemistry of fluorine in group (VII) is different from the rest of the members in the group because;

- Fluorine atom has the smallest atomic radius.
- Fluorine atom has the highest electronegativity value.
- Fluorine molecule the lowest bond dissociation energy.
- Fluorine has the highest electrode potential.
- Fluorine lacks the d-orbital.

Some of the chemical properties in which fluorine differs the rest of the group (vii) elements include;

- Fluoride reacts with water to form hydrogen fluoride and oxygen gas.

  $2F_2 (g) + 2H_2O (l) \longrightarrow 4HF (aq) + O_2(g)$
- Fluoride reacts with carbon to form carbon tetra fluoride

  $2F_2 (g) + C (s) \longrightarrow CF_4 (l)$
- Fluorine can react with cold and dilute sodium hydroxide solution to form sodium fluoride, oxygen difluoride and water.

  $2NaOH (aq) + 2F_2 (g) \longrightarrow 2NaF (aq) + F_2O (g) + H_2O (l)$
- Fluorine reacts with hot and concentrated sodium hydroxide solution to form sodium fluoride, water and oxygen gas.

  $4NaOH (aq) + 2F_2 (g) \longrightarrow 4NaF (aq) + 2H_2O (l) + O_2 (g)$

Diagonal Relationship

This is the similarity in chemical properties between elements in period two and period three of the periodic table that are diagonally adjacent to each other in neighboring groups.

The following pairs of elements in the periodic table show a diagonal relationship;

- Lithium and magnesium
- Beryllium and Aluminum
- Boron and silicon

Reasons for diagonal relationship / why do they do so;

- Have similar charge density.
- Their electronegatives are the same.
- Their standard electrode potentials are similar.

The following chemical properties shows how lithium resembles magnesium and how lithium differs from other group(I) elements:

- ✓ Form ionic nitrides, Li_3N and Mg_3N_2 on heating in nitrogen gas.

  $6 Li (s) + N_2 (g) \longrightarrow 2 Li_3 N (s)$

  $3Mg (s) + N_2 (g) \longrightarrow Mg_3N_2 (s)$
- ✓ Form normal oxides on heating / burring in air

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

  $4 Li (s) + O_2 (g) \longrightarrow 2 Li_2O (s)$
- ✓ Their hydrogen carbonates exists in only solution state.
- ✓ Their carbonates decompose to a metal oxide and carbon dioxide on heating.

  $Li_2CO_3 (s) \longrightarrow Li_2O (s) + CO_2 (g)$

  $MgCO_3 (s) \longrightarrow MgO (s) + CO_2 (g)$
- ✓ Their hydroxides decompose on heating to form metal oxide and water.

  $Mg (OH)_2 (s) \longrightarrow MgO (s) + H_2O (l)$
- ✓ Their nitrates decompose on heating to form metal oxide nitrogen dioxide and oxygen gas.

  $2Mg (NO_3)_2 (s) \longrightarrow 2MgO (s) + 4NO_2 (g) + O_2 (g)$
- ✓ Phosphates, carbonates, fluorides and hydroxides are slightly soluble in water.
- ✓ Halides of magnesium and lithium (except their fluorides) are soluble in organic solvents.
- ✓ Their sulphates do not form alums.

The following shows how beryllium resembles aluminium and how beryllium differs from the other group (ii) elements:

- ✓ Their carbides react with water to form methane gas.

  $Be_2C(s) +4H_2O(l) \longrightarrow 2Be(OH)_2(s) +CH_4(g)$

  $Al_4C_3(s) +12H_2O(l) \longrightarrow 4Al(OH)_2(s) +3CH_4(g)$

- ✓ Their compounds are mainly covalent.
- ✓ Both react with hot concentrated alkalis.

$Be\ (s) + 2OH^- + 2\ H_2O\ (l) \longrightarrow Be(OH)_4^{2-}\ (aq) + H_2\ (g).$

$2Al\ (s) + 2OH^- + 6\ H_2O\ (l) \longrightarrow 2Al(OH)_4^-\ (aq) + 3H_2\ (g)$

- ✓ Their chlorides are easily hydrolyzed in water / fumes in moist air.
- ✓ Their oxides are amphoteric.

$BeO\ (s) + 2OH^-\ (aq) + H_2O\ (l) \longrightarrow Be(OH)_4^{2-}\ (aq)$

$Al_2O_3\ (s) + 2OH^-\ (aq) + 3H_2O\ (l) \longrightarrow 2Al(OH)_4^-\ (aq)$

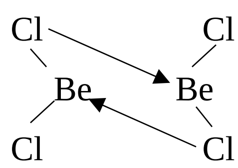
- ✓ Their sulphates are soluble in water.
- ✓ Their hydroxides are amphoteric

$Be(OH)_2\ (s) + 2OH^-\ (aq) \longrightarrow Be(OH)_4^{2-}\ (aq)$

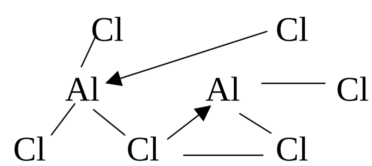
$Al(OH)_3\ (s) + OH^-\ (aq) \longrightarrow Al(OH)_4^-\ (aq)$

- ✓ Their chlorides form dimers in vapor phase

Be_2Cl_4



Al_2Cl_6



- ✓ Their fused chlorides have low boiling point and conductivity.
- ✓ Both metals are passive to nitric acid.

Reasons for diagonal Relationship:

- Have the same charge density that is say ionic charge / ionic radius are very similar.
- Their electro negatives are the same.
- Their electrode potentials are the same.

The ways in which boron resembles silicon and differs from aluminum:

- Both boron and silicon form covalent compounds.
- Both boron and silicon burn in air to form acidic oxides.

$Si + O_2 (g) \longrightarrow SiO_2$

$4B + 3O_2 \longrightarrow 2Bi_2O_3 (s)$

- Their chlorides are hydrolyzed in water to form acids.

$SiCl_4 (l) + 2H_2O (l) \longrightarrow SiO_2 (s) + 4HCl$

$BCl_3 (l) + 3H_2O (l) \longrightarrow H_3BO_3 (aq) + 3HCl (aq)$

- Both form volatile hydrides which easily ignite in air.

- Both react with magnesium and calcium when heated to form borides and silicide.

$2B + Mg \longrightarrow MgB_2$
$6B + Ca \longrightarrow CaB_6$
$2Mg + Si \longrightarrow Mg_2Si$

- When the oxides solidity they form glasses.
- Fluorides are also easily hydrolyzed.

$BF_3 + 3H_2O \longrightarrow 3HBF_4 + B (OH)_3$

$SiF_4 + 4 H_2O \longrightarrow 2H_2SiF_6 + Si (OH)_4$

Group (II) Elements / Alkaline earth metals:

Element	**Symbol**	**Atomic No.**	**Atomic radius**	**M.P °C**	**B.P °C**	**Electronegativity**	**Electronic configuration**
Beryllium	Be	4	0.89	1280	2900	1.47	
Magnesium	Mg	12	1.36	650	1105	1.23	
Calcium	Ca	20	1.74	850	1440	1.04	
Strontium	Sr	38	1.91	770	1370	0.99	
Barium	Ba	56	1.98	720	1140	0.97	

Group (II) elements have a general electronic configuration of $\underline{ns^2}$

The melting point, boiling point and hardness decreases down the group. This is because the atomic radius increases down the group resulting into corresponding decrease in bond strength. The melting point of magnesium is lower than that of calcium because calcium has a face-centered cubic; and magnesium has hexagonal closed-packed.

Chemical properties of group (II) elements:-

1. Reaction with Air.

   They burn brilliantly when heated in air, forming the oxide and a nitride.

   $2Be (s) + O_2 (g) \longrightarrow 2BeO (s)$

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

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

2. Reaction with water.

   Beryllium does not react with water at any condition. Magnesium reacts slowly with cold water to form magnesium hydroxide and hydrogen gas.

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

   Heated magnesium catches fire in steam to form magnesium oxide and hydrogen gas.

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

The rest of the group (II) elements that is strontium, barium and calcium reacts with cold water to form a hydroxide and hydrogen gas.

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

(Y = Ca, Sr, Ba)

Their reactivity with water increases down the group due to the increasing solubility of the hydroxides formed, and due to increasing negativity value of the standard electrode potential.

3. Reaction with acids.

(a)Hydrochloric acid.

They all react with hydrochloric acid on warming to form their respective chlorides and hydrogen gas.

$M (S) + 2H^+ (aq) \longrightarrow M^{2+} (aq) + H_2 (g)$

(M = Be, Ca, Mg, Sr, Ba)

e.g

$Be (s) + 2H^+ (aq) \longrightarrow Be^{2+} (aq) + H_2 (g)$

- **Sulphuric acid**

Dilute sulphuric acid. They all react with warm dilute sulphuric acid to form their respective sulphates and hydrogen gas.

$M (s) + H_2SO_4 (aq) \longrightarrow MSO_4 (aqors) + H_2 (g)$

(M = Be, Ba, Mg, Ca, Sr)

Their reactivity with warm dilute sulphuric acid decreases down the group due to the decreasing solubility of the sulphate formed.

Hot and concentrated sulphuric acid

They all react with hot concentrated sulphuric acid to form a sulphate, sulphur dioxide and water.

$M (s) + 2H_2SO_4 (l) \longrightarrow MSO_4 (aq) + SO_2 (g) + 2H_2O (l).$

Their reactivity with hot concentrated sulphuric acid reduces from Beryllium to barium due to the decreasing solubility of the sulphate formed.

- **Nitric acid**

Beryllium is rendered passive to nitric acid. Magnesium reacts with very dilute nitric acid to form magnesium nitrate and hydrogen gas.

$Mg (s) + 2HNO_3 (aq) \longrightarrow Mg (NO_3)_2 (aq) + H_2 (g)$.

Magnesium reacts with concentrated nitric acid to form oxides of nitrogen, magnesium nitrate and water.

$Mg (s) + 4HNO_3 (aq) \longrightarrow Mg (NO_3)_2 (aq) + 2NO_2 (g) + 2H_2O (l)$

Or

$3Mg (s) + 8HNO_3 (aq) \longrightarrow 3Mg (NO_3)_2 (aq) + 2NO (g) + 4H_2O (l)$

4. Reaction with alkalis.

Only beryllium amongst group (ii) elements the react with alkalis to form beryllate and hydrogen gas.

$Be (s) + 2OH^- (aq) + 2H_2O (l) \longrightarrow Be (OH)_4^{2-} (aq) + H_2 (g)$

5. Reaction with non-metals:

They combine directly at high temperatures (when heated) with the highly electronegative non-metals oxygen, the halogens, nitrogen, sulphur and phosphorus and with carbon. All compounds formed are ionic.

For example;

$Mg (s) + S (s) \longrightarrow MgS (s)$

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

$Sr (s) + Cl_2 (g) \longrightarrow SrCl_2 (s)$

$3Mg (s) + N_2 (g) \longrightarrow Mg_3Ns (s)$

$6Mg (s) + P_4 (s) \longrightarrow 2Mg_3P_2 (s)$

$2Be (s) + C (s) \longrightarrow Be_2C (s)$

$Ca (s) + 2C (s) \longrightarrow CaC_2 (s)$

Calcium, strontium and barium can react / combine with hydrogen gas when heated.

$Ca (s) + H_2 (g) \longrightarrow CaH_2 (s)$

$Sr (s) + H_2 (g) \longrightarrow SrH_2 (s)$

$Ba (s) + H_2 (g) \longrightarrow BaH_2 (s)$

6. Reaction with alkyl halide.

Magnesium can react with alkyl halides to form Grignard reagents in solution.

For example;

$Mg (s) + CH_3I \longrightarrow CH_3MgI$ (methyl magnesium iodide)

Compounds of group (II) elements:

(a) Hydrides.

All form hydrides of general formula MH_2

(Where M = Be, Ca, Mg, Sr, Ba)

They are prepared by heating the metal in dry hydrogen gas.

$$M (s) + H_2 (g) \longrightarrow MH_2 (s)$$

Their reactivity with hydrogen gas reduces down the group.

The hydrides react with water to form hydrogen gas and a hydroxide.

$$MH_2 (s) + 2H_20 (l) \longrightarrow M(OH)_2 (aq) + H_2 (g)$$

(b) Oxides of group (II) elements.

They can be represented by a general formula, MO.
They are prepared by heating the carbonates.

$$MCO_3 (s) \longrightarrow MO (s) + CO_2 (g)$$

Or

By heating their nitrates.

$$2M (NO_3)_2 \longrightarrow 2MO + 4 NO_2 (g) + O_2 (g)$$

Barium oxide can be also be prepared by heating a mixture of barium carbonate and carbon.

$$BaCO_3 (s) + C \longrightarrow BaO (s) + 2CO (g).$$

They can also form peroxides e.g. BaO_2, SrO_2.
The peroxides decompose on heating to form monoxide and oxygen gas.

$$2BaO_2 (s) \longrightarrow 2BaO (s) + O_2 (g)$$

The peroxides dissolve in a cids to form a salt and hydrogen peroxide.

$$BaO_2 (s) + H_2 SO_4 \longrightarrow BaSO_4 + H_2O_2 (aq)$$

$$BaO_2 (s) + 2HCl \longrightarrow BaCl_2 + H_2O_2.$$

All the oxides are basic except beryllium oxide which is amphoteric.

$$BeO (s) + 2OH^- (aq) + H_2O (l) \longrightarrow Be (OH)_4^{2-} (aq)$$

Beryllium oxide does not dissolve in water, the other oxides dissolve in water to form hydroxides, their reactivity with water increases from magnesium oxide to barium oxide through calcium oxide and strontium oxide due to the increasing solubility of the hydroxide formed.

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

$BaO\ (s) + H_2O\ (l) \longrightarrow Ba(OH)_2\ (aq)$

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

They react with acidic oxides to form salts.
For example;

$6CaO + P_4O_{10} \longrightarrow 2Ca_3(PO_4)_2$

$MgO + SO_3\ (g) \longrightarrow MgSO_4\ (s)$

(c) Hydroxides.

Hydroxide	$Be(OH)_2$	$Mg(OH)_2$	$Ca(OH)_2$	$Sr(OH)_2$	$Ba(OH)_2$
Solubility	Insoluble	0.002	0.15	0.9	4

The solubility of the hydroxides increases from $Be(OH)_2$ to $Ba(OH)_2$ through $Mg(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$.

I.e. $Be(OH)_2 < Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$ due to increase in cationic radius, causes lattice energy and hydration energy decreases but lattice energy decreases more than hydration energy making the solubility to increase.

Only beryllium hydroxide is an amphoteric oxide.

$Be(OH)_2\ (s) + 2OH^-(aq) \longrightarrow Be(OH)_4^{2-}\ (aq)$

The rest of the hydroxides are basic.

All hydroxides decompose on heating to form oxide and water

$M(OH)_2\ (s) \longrightarrow MO\ (s) + H_2O\ (l)$

(d) **Carbides,** calcium dicarbide CaC_2,

calcium dicarbide is prepared by heating carbon (coke) and calcium oxide at 2000°C.

$CaO\ (s) + C \longrightarrow CaC_2\ (s) + CO\ (g)$

Calcium dicarbide dissolves in water to form ethyne

$CaC_2\ (s) + 2H_2O\ (l) \longrightarrow C_2H_2\ (g) + Ca(OH)_2\ (aq)$

Beryllium carbide dissolves in water to form methane

$Be_2C(s) + 4H_2O(l) \longrightarrow 2Be(OH)_2(s) + CH_4(g)$

(e) Sulphate.

The order of the solubility of sulphates is;

$BeSO_4$ > $MgSO_4$ > CaSO4 > $SrSO_4$ > $BaSO_4$ due to increasing cationic radius, lattice energy and hydration energy decreases, but hydration energy decreases more than lattice energy, making the solubility to reduce.

(f) Carbonates.

The order of decomposition of the carbonates is;

$BeCO_3 < MgCO_3 < CaCO_3 < SrCO_3 < BaCO_3$

This is mainly due to increase in cationic radius which results into increase in the ionic character of the carbonates, making the thermostability of the carbonates to increase.
They are prepared by precipitation.

$M^{2+} (aq) + CO^{2}{}_{3}{}^{-} \longrightarrow MCO_3 (s)$

(g) Calcium phosphate, $Ca_3 (PO_4)_2$

It is a rock used in the preparation of fertilizer – super phosphate made by reacting sulphuric acid and calcium phosphate.

$Ca_3 (PO_4)_2 + 2H_2SO_4 (aq) \longrightarrow Ca(H_2PO_4)_2 (s) + 2CaSO_4 (s)$