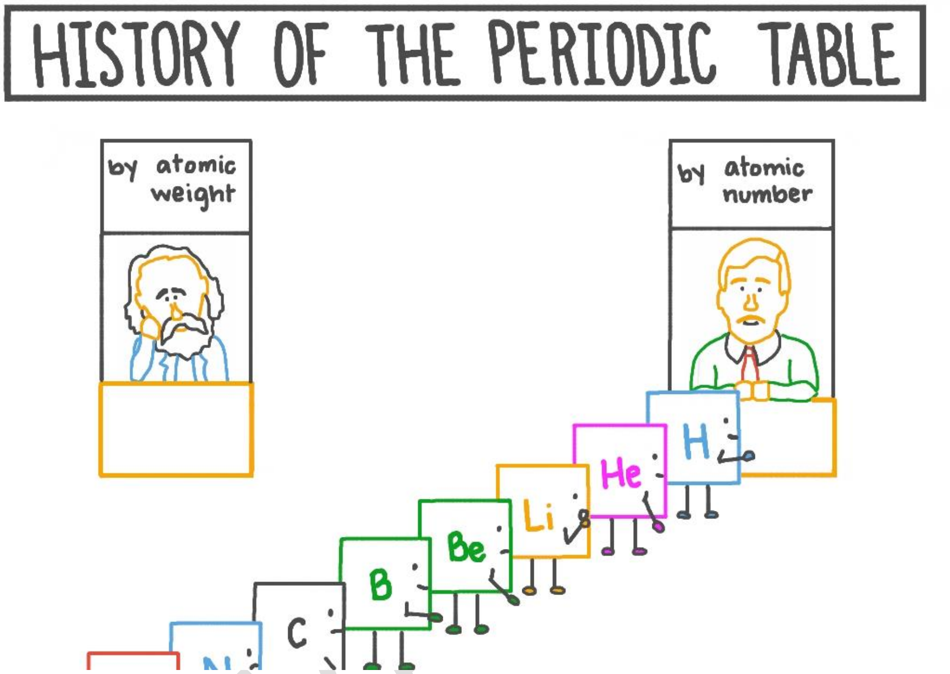
PERIODICITY I

Competence: ***The learner analyses the trends and periodic properties of elements, to explain and predict the reactivity and properties of elements in the Periodic Table.***

THE PERIODIC TABLE

Historical Development:

The development of the periodic table involved key contributions from Dmitri Mendeleev and Henry Moseley. However, the there are some earlier versions which did not impact much on the development of the Periodic Table.



1. Early Attempts Before Dmitri Mendeleev.

(a) Dobereiner's Triads (1829):

(i) Grouped elements in threes based on similar properties.

(ii) Middle element's atomic mass ≈ average of the other two.

Example: Ca, Sr, Ba.

(b) Newlands' Law of Octaves (1864):

(i) Arranged elements by increasing atomic mass.

(ii) Every 8th element had similar properties.

(iii) Failed for elements beyond calcium.

2. Dmitri Mendeleev's Periodic Table (1869).

(a) Arrangement Principle:

Mendeleev arranged elements in order of increasing atomic mass, noting that elements with similar chemical properties appeared at regular intervals into vertical columns.

(b) Key Achievements:

He grouped elements with similar properties into vertical columns (groups).

(i) Prediction of Missing Elements.

He left gaps in the table for elements not yet discovered (e.g., gallium, scandium, germanium), and accurately predicted their properties based on the patterns he observed.

(ii) Correction of Atomic Masses.

He prioritized chemical behavior over strict atomic mass order, sometimes placing elements slightly out of order based on their chemical properties.

For example, he adjusted by swapping the positions of iodine and tellurium based on (to fit the pattern of) their chemical similarities to other group VII elements (halogens) and group VI elements, respectively even if mass order seemed “wrong”.

(iii) Periodic Law (Mendeleev’s version):

"The properties of the elements are a periodic function of their atomic masses."

(c) Significance of Mendeleev’s version:

Mendeleev's table was the first widely recognized and useful organization of the elements, and it demonstrated the periodic law (that properties of elements are a periodic function of their atomic weights).

(d) Limitations of Mendeleev’s Table:

(i) Position of some elements (like argon and potassium) couldn’t be fully explained.

(ii) Did not account for isotopes.

(iii) Could not explain why periodicity occurred.

3. Henry Moseley’s Atomic Concept (1913)

(a) Discovery:

(i) Moseley discovered that each element has a unique atomic number, which corresponds to the number of protons in its nucleus.

(ii) He used X-ray spectra to determine the nuclear charge (atomic number) of elements.

(b) Impact on Periodic Table:

(i) Arrangement.

He arranged the elements in the periodic table by increasing atomic number, rather than atomic mass.

(ii) Significance.

Moseley's work provided a more accurate and scientific basis for the periodic table, resolving inconsistencies in Mendeleev's table and explaining why certain elements were placed where they were.

Argon (Z=18) correctly placed before potassium (Z=19) despite having a higher atomic mass.

In essence, Mendeleev provided the initial framework, while Moseley provided the crucial refinement that led to the modern periodic table.

Mendeleev organized elements by atomic mass and chemical properties, leaving gaps for undiscovered elements and even predicting their properties. Moseley later refined this by arranging elements by atomic number (number of protons), which accurately reflected the elements' chemical behavior and corrected inconsistencies in Mendeleev's arrangement.

(iii) Led to the Modern Periodic Law.

It also confirmed the periodic law, but with atomic number as the key factor.

"The properties of elements are a periodic function of their atomic numbers."

(c) Modern Arrangement:

(i) The modern Periodic Table consists of 7 periods (horizontal rows) and 18 groups (vertical columns).

(ii) It accounts for isotopes naturally.

4. Key Differences Between Mendeleev’s & Moseley’s Approaches.

Feature	Mendeleev (1869)	Moseley (1913)
Basis of arrangement	*Atomic mass*	*Atomic number*
Law stated	*Periodic properties depend on atomic mass*	*Periodic properties depend on atomic number*
Problem with anomalies	*Some elements out of order by mass*	*No anomalies in order*
Missing elements	*Left gaps and predicted their properties*	*Arrangement confirmed predictions*

5. Summary Diagram for the Historical Flow.

Dobereiner → Newlands → Mendeleev (atomic mass) → Moseley (atomic number) → Modern Periodic Table.

The Periodic Classification of Elements

Classification of all chemical elements is based on the Modern Periodic Law which states that *"the properties of elements are periodic functions of their atomic numbers"*.

In other words, if elements are arranged in order of increasing atomic number, then elements with similar chemical properties occur at regular intervals and in the same group (vertical column).

Elements are classified into blocks according to the type of atomic orbitals or sub-energy level in which the last electron is added in accordance to Aufbau's principle. These blocks are:

- The s-block. - The p-block. - The d-block. - The f-block.

(a) **The s-block elements:**

These are elements whose last electron is filled in the s-orbital (or s-sub energy levels) of the valence shell. The elements in this block have a general outer shell electronic configuration of ns^1 (Group I) and ns^2 (Group II). where n is the principal quantum number (quantum number of the outermost energy level).

(b) **The p-block elements:**

These are elements whose last electron is filled in the p-orbitals after the s-orbital in the outermost energy level (valence shell) has been filled.

The elements in this block have a general outermost shell electronic configuration ns^2np^x where x = 1,2, 3, 4, 5, 6

The p-block elements therefore comprise of elements in Groups III, IV, V, VI, VII and VIII.

(c) **The d-block elements:**

These are elements whose last electron is filled in the d-orbitals of the **penultimate shell** (the inner shell that follows the valence shell). The shell which is before the valence shell is called penultimate shell. If 'n' is the valence shell then (n-1) is the penultimate shell. Transition elements contain unpaired electrons in penultimate shells.

These elements occur between the s-block and the p-blocks and form 4 series of elements i.e.

(i) Scandium (Sc) with atomic number 21 to Zinc (Zn) with atomic number 30 in period 4.

(ii) Yttrium (Y) with atomic number 39 to Cadmium (Cd) with atomic number 48 in period 5.

(iii) Hafnium (Hf) with atomic number 72 to Mercury (Hg) with atomic number 80 in period 6.

(iv) Rutherfordium (Rf) with atomic number 104 to Copernicium (Cn) with atomic number 112 in period 7.

(d) **The f-block elements:**

These are elements in which the last electron is filled in the f-orbitals of the **antepenultimate energy level** (3rd last energy level) in the atoms across the period.

There are two series of f-block elements namely:
(i) The lanthanides (lanthanoids) consisting of elements Lanthanum (La) to lutetium (Lu).
(ii) The actinides (actinoids) consisting of elements Actinum (Ac) to Lawrencium (Lr).
N.B: Ante-penultimate shell is the n-2 shell i.e. 2 shells before the valence shell. The outermost shell is the valence shell, the one before it is penultimate shell and the one before it is called ante-penultimate shell. Hence the ante-penultimate shell is the shell third from the last.

Position of Elements and Their Reactivity

The position of an element in the periodic table dictates its chemical behavior due to the predictable trends in properties based on electron configuration.

1. Reading the Element's Position:

To predict chemical behaviour, three key things are first noted from its place on the Periodic Table:
(a) Group (vertical column). This tells the number of valence electrons and hence similar chemical reactivity within a group.
(b) Period (horizontal row). This tells the number of energy levels (electron shells) and hence influences atomic radius (size) and reactivity. Moving across a period (horizontal row), properties change gradually, reflecting the changing number of valence electrons and their influence in chemical reactions.
(c) Block (s, p, d, f). This indicates the type of orbital being filled, which affects metallic/non-metallic character.

2. Some Key or Major Periodic Trends and Their Influence on Chemical Behaviour:

By understanding these trends, one can predict the general chemical behavior of an element based on its position in the periodic table.
(a) Valence Electrons. The number of valence electrons (electrons in the outermost shell) determines how an atom will bond with other atoms. Elements in the same group have the same number of valence electrons, leading to similar bonding patterns and chemical properties.
For example, all halogens (Group VII) have 7 valence electrons and readily gain one electron to achieve a stable octet, forming negative ions.
(b) Metallic Character. Metallic character generally decreases from left to right across a period and increases down a group. Metals tend to lose electrons to form positive ions, while nonmetals tend to gain electrons. The left side of the periodic table is dominated by metals, while the right side is dominated by nonmetals.
(c) Electronegativity. Electronegativity, the ability of an atom to attract electrons in a covalent bond, generally, increases across a period and decreases down a group. Fluorine (F) is the most electronegative element. Higher electronegativity indicates a greater tendency to gain electrons.
(d) Atomic Radius. Atomic radius generally increases down a group and decreases across a period. As you move down a group, the number of energy levels increases, leading to a larger atomic radius. Moving across a period, the increasing nuclear charge pulls the electrons closer, resulting in a smaller atomic radius.
(e) Ionization Energy. Ionization energy, the energy required to remove an electron, generally increases across a period and decreases down a group. Elements with high ionization energies tend to be less reactive and more likely to form negative ions. Elements with low ionization energies tend to be more reactive and readily form positive ions.

Trend	Across a Period (→)	Down a Group (↓)	Chemical Behaviour Impact
Atomic Radius	Decreases (more protons	Increases (more	Small atoms hold electrons tightly; large

Trend	Across a Period (→)	Down a Group (↓)	Chemical Behaviour Impact
	pulling electrons closer)	shells)	atoms lose them more easily.
Ionization Energy	Increases (harder to remove electrons)	Decreases (outer electrons farther from nucleus)	Low ionization energy enables metals to lose electrons to form cations; High ionization energy of nonmetals enables them to gain electrons and hence form anions.
Electronegativity	Increases (nucleus attracts electrons more strongly)	Decreases	High electronegativity enables the element to be strong oxidizing agents; Low electronegativity enables the element to be a strong reducing agent.
Metallic Character	Decreases	Increases	Metals can conduct, form basic oxides; Nonmetals are insulators, form acidic oxides.

3. Predicting Chemical Behaviour Using Position.

Reactivity is closely linked to the tendency to gain or lose electrons.

Highly reactive metals tend to have low ionization energies and readily lose electrons to form positive ions.

Highly reactive nonmetals tend to have high electronegativity and readily gain electrons to form negative ions.

The reactivity of alkali metals (Group I) increases as you move down the group, while the reactivity of halogens (Group VII) generally decreases as you move down the group.

(a) Metals (left and center of table). Generally, they:
- have low ionization energy, low electronegativity.
- tend to lose electrons and form positive ions (cations).
- react with nonmetals to form ionic compounds.
- react with acids to produce hydrogen gas. etc.

(b) Nonmetals (upper right). Generally, they:
- have high ionization energy, high electronegativity.
- tend to gain electrons and form negative ions (anions).
- often form covalent bonds with other nonmetals. etc.

(c) Noble Gases (Group VIII). Generally, they are very stable and inert due to having a completely filled valence shell. However, heavier ones (like Xe) can form compounds under special conditions.

(d) Transition Metals (d-block). They:
- have variable oxidation states.
- have catalytic properties.
- are paramagnetic. etc.
- often form coloured compounds.
- form complex ions.

4. Examples of Predictions.

(a) Sodium in Group I and Period 3.

It has a large atomic radius, low ionisation energy, low electronegativity hence it will lose 1 electron easily. This is observed when it reacts violently with water to form sodium hydroxide and hydrogen gas. etc

(b) Fluorine in Group VII (halogens) and Period 2.

It has a small atomic radius, very high electronegativity and will gain 1 electron strongly.

It is a very reactive oxidizing agent.   etc.
(c) Magnesium in Group II (alkaline earth metals) and Period 3.
It has a low ionisation energy hence it loses 2 electrons to form magnesium ions, Mg^{2+}.
Readily reacts with acids and burns in air to form magnesium oxide, MgO. etc.

✅ **The Bottom Line:**

Knowing where an element sits within the periodic table, enables the ability to predict whether it will lose or gain electrons, how strongly it will react, what kind of compounds it will form, and even whether it will be coloured, acidic, basic, or inert.

PERIODICITY OF PHYSICAL PROPERTIES OF ELEMENTS

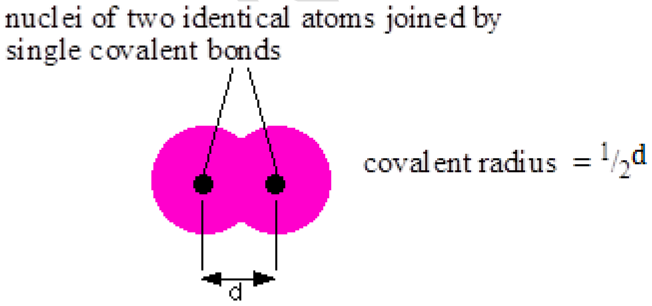
(VARIATION OF PHYSICAL PROPERTIES IN THE PERIODIC TABLE)

Periodicity is *the repeating pattern of both physical and chemical properties of elements as a function of their atomic numbers*.
In other words, a plot of graph of a physical property against atomic number of elements consists of a series of peaks (maxima) and troughs (minima) clearly showing that the property under investigation or consideration is the periodic function of the atomic number of elements.
Periodic trends or patterns are obtained by physical properties of elements such as **atomic radius, ionization energy, electron affinity, electronegativity, electropositivity, melting point, boiling point, atomization energy, standard electrode potential, hydrated radius,** etc.

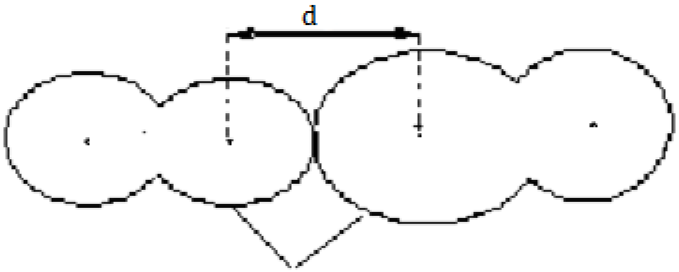
1. ATOMIC RADIUS

Atomic radius is half the distance between the nuclei (inter nuclear distance) of two identical atoms in any bonding situation or in a given bonding situation.
Alternatively, atomic radius is the distance of closest approach to the nucleus of another similar atom in a given bonding situation.
Unlike a ball, an atom doesn't have a fixed radius. The radius of an atom can only be found by halving the measured distance between the nuclei of two bonding or touching atoms.
Units for measuring atomic radius can be nanometers (nm) or angstroms (Å).
There are three types of **atomic radii.**

(a) **Covalent radius.**
This is half the distance between nuclei of two identical atoms holding together with a single covalent bond.



(b) **Van der Waal's Radius.**
This is half the distance between the nuclei of two identical atoms in adjacent molecules held together by Van der Waal's forces in the solid state.
The illustration below is the diatomic molecule of iodine held together by Van der Waal's force of attraction.



Van der Waal's radius = $\frac{d}{2}$

Two molecules held together by Van der Waal's forces in the solid state. eg. iodine molecules.

(c) Metallic radius.

This is half the distance between the nuclei of two identical metal ions in a metallic crystal.

Factors Affecting Atomic Radius

These are concerned with the measure of attraction between the nucleus and the electron(s) in the outermost energy level.

(a) **Nuclear charge.** Nuclear charge is the total charge in the nucleus of an atom or ion for all the protons. Thus, charge on the nucleus is positive which depends on the number of protons in the nucleus. The higher the proton number, the higher the nuclear charge.

Trend: The higher the nuclear charge the smaller the atomic radius.

Explanation: When the nuclear charge increases due to increase in the number of protons, there is a stronger nuclear attraction for the outermost energy level electrons in the atom hence leading to a reduction in the atomic radius.

Similarly, the lower the proton number, the lower the nuclear charge and the larger the atomic radius. This is because, when the nuclear charge is lower due decrease in the number of protons, there is a weaker nuclear attraction for the outermost energy level electrons hence leading to an increase in the atomic radius.

(b) **Screening (Shielding) effect.** Screening effect is the mutual repulsion that takes place between electrons particularly in the inner energy levels with those in the outermost energy level.

Trend: The higher the screening effect, the larger the atomic radius.

Explanation: When the screening effect increases due to an increase in the number of electrons, the electrons in the outermost energy level move to a distance further away from the nucleus. Thus, the outermost electrons become weakly attracted by the positive charge on the nucleus resulting into an increase in atomic radius.

Similarly, the lower the screening effect, the smaller the atomic radius. This is because when the screening effect is lower due to a decrease in the number of electrons, the electrons in the outermost energy level move to a distance closer to the nucleus. Thus, the outermost electrons become strongly attracted by the positive charge on the nucleus resulting into a decrease in atomic radius.

Note: (i) Screening effect significantly increases with increase in number of inner energy levels that are completely filled with electrons. However, screening effect slightly increases (remains almost constant) as the number of electrons increase within the same energy level.

(ii) Within a given energy level, the screening/shielding effect of the inner electrons decreases in the order; s > p > d > f

(iii) The *effective nuclear charge* (Z_{eff}) is the net (overall) positive charge that an electron in the outermost energy level experiences from the nucleus after considering the influence of the nuclear charge and the screening effect.

i.e. It is the actual amount of positive charge experienced by an electron in a polyelectronic atom.

The effective nuclear charge is given by the expression below:

$Z_{eff} = Z - S$ where Z = nuclear charge S = screening effect

e.g. In the lithium atom, Z = 3 and S = 1.7 hence $Z_{eff} = 3 - 1.7 = 1.3$

In the sodium atom, Z = 11 and S = 10 hence $Z_{eff} = 11 - 10 = 1$

Thus, effective nuclear charge is always lower than the nuclear charge.

On analysis of trends across the period and down the group; when the increase in the nuclear charge is greater than (outweighs/exceeds) the increase in the screening effect, then the effective nuclear charge increases. However, when the increase in screening effect is greater than the increase in the nuclear charge then the effective nuclear charge decreases.

Variation of atomic radius across periods 2 and 3 of the Periodic Table

Period 2

Element	Li	Be	B	C	N	O	F
Atomic radius (nm)	0.123	0.089	0.080	0.077	0.074	0.073	0.072
Atomic number	3	4	5	6	7	8	9

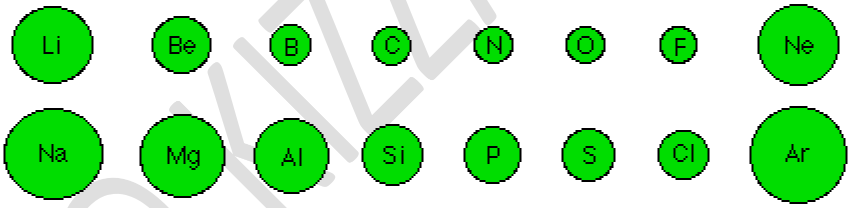
Period 3

Element	Na	Mg	Al	Si	P	S	Cl
Atomic radius (nm)	0.157	0.136	0.125	0.117	0.110	0.104	0.099
Atomic number	11	12	13	14	15	16	17

Task: Plot separate graphs of atomic radius against atomic number of the Period 2 and Period 3 elements.

Trend: Atomic radius decreases across elements of period 2 and 3 in the Periodic Table.

Explanation: This is because on moving across the period from one element to the next, there is addition of an electron to the same energy level which leads to a slight increase in the screening effect. There is also addition of a proton to the nucleus which leads to an increase in the nuclear charge. However, the increase in the nuclear charge is greater than the increase in screening effect hence the effective nuclear charge increases. Thus, the nuclear attraction for the outermost electrons becomes stronger leading to a decrease in the atomic radius.



Note that: The atomic radius of noble gases is bigger due to its Van der Waal's radius.

Variation of atomic radius down a group of elements of the Periodic Table

Group I

Element	Li	Na	K	Rb	Cs
Atomic radius (nm)	0.123	0.157	0.203	0.216	0.235

Group II

Element	Be	Mg	Ca	Sr	Ba
Atomic radius (nm)	0.089	0.136	0.174	0.194	0.198

Trend: Atomic radius increases down the group of elements in the periodic table.

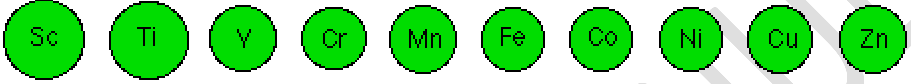
Explanation: This is because on moving down the group from one element to the next, there is addition of a new **inner** energy level completely filled with electrons which leads to a large increase in the screening effect.

There is also addition of protons to the nucleus which leads to an increase in the nuclear charge.
However, the increase in the screening effect is greater than the increase in nuclear charge hence the effective nuclear charge decreases. Thus, the nuclear attraction for the outermost electrons becomes weaker leading to an increase in the atomic radius.

Trend of atomic radius among d-block elements (Transition metals) in Period 4

Atomic radii of transition elements are smaller as compared to group I and II elements because the electrons being added to the inner 3d sub-energy level of the third energy level and are nearer the nucleus than the electrons in the outermost 4s orbitals in the fourth energy level.

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic radius (nm)	0.144	0.132	0.122	0.117	0.117	0.116	0.116	0.115	0.117	0.125
Atomic number	21	22	23	24	25	26	27	28	29	30



Atomic radii remain almost constant and when they vary, they do not follow a definite pattern.

N.B: - The decrease in atomic radius is small as compared to the s and p block elements because of the screening effect caused by electrons of the inner (n-1)d subshell on the outermost shell.

- Screening effect becomes significant or effective when electrons are being added to the inner orbitals than outer orbitals.

- The radius of transition metals is determined by the outermost 4s orbital electrons in the fourth energy level.

Trend: There is a slight general decrease in atomic radii from scandium to nickel as the atomic number increases but slightly increases up to zinc.

Explanation: This is because on moving across the period from one element to the next, there is addition of an electron to the same inner 3d sub energy level (penultimate shell) which leads to a slight increase in the screening effect. There is also addition of a proton to the nucleus which leads to an increase in the nuclear charge. The increase in nuclear charge is slightly greater than the increase in the screening effect due to the electrons in partly filled inner 3d orbitals and this leads to a slight increase in the nuclear attraction for the outermost 4s electrons hence a slight decrease in the atomic radius from scandium to nickel.
However, as the inner 3d-orbitals are completely filled, it results in change of patterns in the nuclear attraction and screening effect. The slight increase in atomic radii from nickel to copper is because the 3d orbitals are completely filled and their electrons more readily screen the outermost 4s orbital electrons. The increase in the screening effect becomes slightly greater than the increase in the nuclear charge.

IONIC RADIUS

This is half the inter nuclear distance between two neighbouring identical ions in contact with each other in a crystal lattice of an ionic compound. Ionic radii are difficult to measure because it varies according to the environment of the ion. i.e the type of oppositely charged ion with which an ion is chemically combined, the co-ordination of the ion (how many ions of same charge are directly bonded to an oppositely charged ion), etc.

(a) **Cationic radius.**

Electropositive elements (e.g. metals) readily form positive ions (cations). The radius of a cation is smaller than the radius of the atom from which it is derived. e.g.

i) The atomic and ionic radii of Na and Na^+ are 0.157nm and 0.095nm respectively.
ii) The atomic and ionic radii of Mg and Mg^{2+} are 0.136nm and 0.065nm respectively.
iii) The atomic and ionic radii of Fe = 1.17Å, Fe^{2+} = 0.76Å, Fe^{3+} = 0.64Å.

Explanation: This is because after loss of an electron or electrons, the number of electrons decreases but the number of protons remains constant. The proton to electron ratio increases and the nuclear attraction of the remaining electrons becomes stronger. Hence the electrons are more strongly attracted to the nucleus resulting into a smaller radius for the cation formed.

There is also reduction in the number of energy levels in some cations when all the valence electrons are lost. Thus, cations contain fewer electrons and fewer energy levels compared to the atoms from which they are obtained.

(b) **Anionic radius.**

Electronegative elements (e.g. nonmetals) form negative ions (anions).

The radius of an anion (negative ion) is larger than that of the atom from which it is derived. e.g. i) The atomic and ionic radii of O and O^{2-} are 0.073nm and 0.134nm respectively.

ii) The atomic and ionic radii of P and P^{3-} are 0.11nm and 0.212nm respectively.

iii) The atomic and ionic radii of Cl = 1.40Å, Cl^- = 1.81Å.

Explanation: This is because after addition of an electron or electrons, the number of electrons increases but the number of protons remains constant. The proton to electron ratio decreases, and there is more repulsion between electrons such that the outermost electron moves to some distance further away from the nucleus. Thus, the nuclear attraction for the outermost electron becomes weaker resulting into a larger radius of the anion formed.

Isoelectronic Series of Ions

This is a group of ions with the same number of electrons and hence have the same electronic configuration. e.g.

(a) **Variation of ionic radii of isoelectronic ions in different periods but with increasing number of protons. e.g. period 2 and Period 3.**

i.e N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, Al^{3+} - with 10 electrons each and an electronic configuration $1s^22s^22p^6$

Trend: The ionic radius decreases across the periods.

Explanation: Across the periods as we move from one ion to the next, a proton is added to the nucleus of an ion. The nuclear charge increases, which easily outweighs the constant screening effect. The effective nuclear charge greatly increases and the nuclear attraction for the electrons becomes stronger hence leading to a decrease in the ionic radii.

Ion	**N^{3-}**	**O^{2-}**	**F^-**		**Na^+**	**Mg^{2+}**	**Al^{3+}**
Number of protons	7	8	9		11	12	13
Electronic Structure	$1s^22s^22p^6$	$1s^22s^22p^6$	$1s^22s^22p^6$		$1s^22s^22p^6$	$1s^22s^22p^6$	$1s^22s^22p^6$
Ionic radius (nm)	0.171	0.140	0.133		0.102	0.072	0.054

Similar reasoning explains why the radius of ions in series with 18 electrons decreases from P^{3-} to Ca^{2+}. i.e. $P^{3-} > S^{2-} > Cl^- > K^+ > Ca^{2+}$

(b) **Variation of ionic radii of ions in same period. e.g. Period 3.**

(i) Na^+, Mg^{2+}, Al^{3+} with 10 electrons each and an electronic configuration; $1s^22s^22p^6$.

(ii) P^{3-}, S^{2-}, Cl^- with 18 electrons each and an electronic configuration; $1s^22s^22p^63s^23p^6$. However, the ions have different proton numbers hence different nuclear charges and therefore different ionic radii.

Ion	**Na^+**	**Mg^{2+}**	**Al^{3+}**	**P^{3-}**	**S^{2-}**	**Cl^-**
Number of protons	11	12	13	15	16	17
Electronic Structure	$1s^22s^22p^6$	$1s^22s^22p^6$	$1s^22s^22p^6$	$1s^22s^22p^63s^23p^6$	$1s^22s^22p^63s^23p^6$	$1s^22s^22p^63s^23p^6$
Ionic radius (nm)	0.102	0.072	0.054	0.212	0.184	0.181

Trend: The ionic radius decreases across the period; however, the anionic radii are larger than the cationic radii.

Explanation: Across the period as we move from one ion to the next, a proton is added to the nucleus of an ion but the electrons remain the same. The nuclear charge increases, which easily outweighs the constant screening effect. The effective nuclear charge greatly increases and the nuclear attraction for the electrons increases. This causes the ionic radii to decrease.

However, the anionic radii are generally larger that the cationic radii due to addition of an extra energy level. Thus, the screening effect of anions is higher than that of cations leading to a smaller increase in the effective nuclear charge as nuclear attraction for the outermost electrons becomes weaker.

2. IONISATION ENERGY (POTENTIAL)

Ionisation energy is the minimum amount of energy required to remove one mole of electrons from one mole of isolated/free gaseous atoms or cations to form one mole of isolated/free gaseous cations.

Alternatively, ionization energy is the minimum amount of energy required to remove an electron from an isolated/free gaseous atom or cation to form a gaseous cation.

i.e. $M_{(g)} - e \rightarrow M^+_{(g)}$ or written as $M_{(g)} \rightarrow M^+_{(g)} + e$ $\Delta H = 1^{st}$ I.E ($kJmol^{-1}$)

1 mole 1 mole 1 mole

$M^+_{(g)} - e \rightarrow M^{2+}_{(g)}$ or written as $M^+_{(g)} \rightarrow M^{2+}_{(g)} + e$ $\Delta H = 2^{nd}$ I.E ($kJmol^{-1}$)

1 mole 1 mole 1 mole

Types of ionization energies

(a) **First ionization energy.**

This is the minimum amount of energy required to remove one mole of electrons from one mole of isolated/free gaseous atoms to form one mole of singly charged or univalent gaseous cations (unipositively charged gaseous ions).

Alternatively, this is the minimum amount of energy required to remove an electron from an isolated/free gaseous atom to form a singly charged or univalent gaseous cation (unipositively charged gaseous ion).

e.g. for sodium; $Na_{(g)} \rightarrow Na^+_{(g)} + e$ $\Delta H = +494\ kJmol^{-1}$

N.B: When tasked to define the term ionisation energy, a candidate is advised to state the definition of first ionisation energy.

(b) **Second ionization energy.**

This is the minimum amount of energy required to remove one mole of electrons from one mole of isolated singly charged gaseous cations to form one mole of doubly/divalent charged gaseous cations.

e.g. for sodium; $Na^+_{(g)} \rightarrow Na^{2+}_{(g)} + e$ $\Delta H = +4563\ kJmol^{-1}$

Other types of ionization energies include 3rd, 4th, 5th etc ionization energies depending on the number of electrons present in the atom.

Factors that affect the ionization energy

These are concerned with the measure of attraction between the nucleus and the outermost electron (the electron in the last energy level).

(a) **Atomic radius**

Trend: The smaller the atomic radius, the higher the ionization energy. Alternatively, the larger the atomic radius, the lower the ionization energy.

Explanation: When the radius of an atom is smaller, then the nuclear attraction for the outermost electron(s) in the atom is stronger. Thus, more energy is required to remove the outermost electron from a gaseous atom of the element, hence having a higher ionization energy.

Alternatively, when the radius of the atom is larger, then the nuclear attraction for the outermost electron is weaker. Thus, less energy is required to remove the outermost electron from a gaseous atom hence having lower ionization energy.

e.g. Aluminium; $Al_{(g)} \rightarrow Al^{+}_{(g)} + e \quad \Delta H = +577\ kJmol^{-1}$

Silicon; $Si_{(g)} \rightarrow Si^{+}_{(g)} + e \quad \Delta H = +786\ kJmol^{-1}$

(b) **Nuclear charge**

Trend: The higher the nuclear charge, the higher the ionization energy. Alternatively, the lower the nuclear charge, the lower the ionization energy.

Explanation: When the nuclear charge of the atom is higher, then the nuclear attraction for the outermost electron is stronger. Thus, more energy is required to remove the outermost electron from a gaseous atom hence having higher ionization energy.

Alternatively, when the nuclear charge is lower, then the nuclear attraction in the outermost is weaker. Thus, less energy is required to remove the outermost electron and hence lower ionization energy.

e.g. Aluminium; $Al_{(g)} \rightarrow Al^{+}_{(g)} + e \quad \Delta H = +577\ kJmol^{-1}$

Silicon; $Si_{(g)} \rightarrow Si^{+}_{(g)} + e \quad \Delta H = +786\ kJmol^{-1}$

(c) **Screening/Shielding effect**

Trend: The higher the screening effect, the lower the ionization energy. The reverse is true.

Explanation: The higher the screening effect, the more the outermost electron moves to a distance further away from the nucleus. Thus, the weaker the nuclear attraction for the outermost electron/electrons and a smaller amount of energy is required to remove the outermost electron from a gaseous atom hence a lower ionization energy. Alternatively, the lower the screening effect, the stronger the nuclear attraction for the outermost electrons. Hence the larger the amount of energy required to remove the outermost electron and the higher the ionization energy value.

e.g. Magnesium; $Mg_{(g)} \rightarrow Mg^{+}_{(g)} + e \quad \Delta H = +737\ kJmol^{-1}$

Calcium; $Ca_{(g)} \rightarrow Ca^{+}_{(g)} + e \quad \Delta H = +590\ kJmol^{-1}$

(d) **The electronic configuration of the atom or ion**

Atoms (or ions) with fully filled or half-filled orbitals (sub energy levels) have higher ionization energies than expected.

This is because fully filled or half-filled orbitals are thermodynamically (energetically) stable while those that are neither half-filled nor fully filled are thermodynamically unstable.

Thus, removal of an electron from orbitals that are half filled or fully filled does not easily take place and a lot of energy is absorbed to effect removal of the electron hence having higher ionization energies than expected.

N.B: Fully filled orbitals are more thermodynamically stable than half-filled orbitals.

$Mg_{(g)} \rightarrow Mg^{+}_{(g)} + e$ $\Delta H = +737\ kJmol^{-1}$ $1s^2 2s^2 2p^6 3s^2$ higher than that for aluminium

$P_{(g)} \rightarrow P^{+}_{(g)} + e$ $\Delta H = +1012\ kJmol^{-1}$ $1s^2 2s^2 2p^6 3s^2 3p^3$ higher than that for sulphur

Element	Na	Mg	Al	Si	P	S	Cl	Ar
1st I.E ($kJmol^{-1}$)	+494	+737	+577	+786	+1012	+999	+1255	+1521

(e) **Penetrating power of orbitals**.

Penetrating power is the tendency of orbitals to approach the nucleus more closely than others.

Trend: The higher the penetrating power, the higher the ionization energy.

Explanation: The higher the penetrating power of orbitals, the stronger the nuclear attraction for the electrons in the orbitals. Thus, the nuclear attraction for the electron is stronger and more energy is required to remove an electron in those orbitals hence having a higher ionization energy.

N.B: Electrons occupy orbitals which have different penetrating powers i.e. different abilities to approach the nucleus. This is because different orbitals have different shapes. The penetrating power of electrons in the different types of orbitals decreases in the order; s > p > d > f.

The s-orbital electrons have highest penetrating power and therefore require the highest amount of energy to be removed while the f-orbital electrons have the least penetrating power and so require the lowest amount of energy to be removed.

(f) **Net charge on an ion within the same period.**

(i) Cations.

Trend: The higher the net charge on the cation, the higher the ionization energy. Alternatively, the lower the net charge on the cation, the lower the ionization energy.

Explanation: When the net charge on a cation is higher, then the proton to electron ratio is also higher. The nuclear attraction for the outermost electron becomes stronger and more energy is required to remove an electron from the cation, hence having higher ionization energy.

Alternatively, when the net charge on the cation is lower, then the proton to electron ratio is also lower. The nuclear attraction for the outermost electron becomes weaker and less energy is required to remove an electron from the cation, hence having lower ionization energy.

$Be^{2+}_{(g)} \rightarrow Be^{3+}_{(g)} + e^-$ 3rd ionization energy = +14905 kJ/mol

$Be^{3+}_{(g)} \rightarrow Be^{4+}_{(g)} + e^-$ 4th ionization energy = +21060 kJ/mol

(ii) Anions.

Trend: The higher the net charge on the anion, the lower the ionization energy. Alternatively, the lower the net charge on the anion, the higher the ionization energy.

Explanation: When the net charge on an anion is higher, then the proton to electron ratio is lower. The nuclear attraction for the outermost electron becomes weaker and less energy is required to remove an electron from the anion, hence having lower ionization energy.

Alternatively, when the net charge on the anion is lower, then the proton to electron ratio is higher. The nuclear attraction for the outermost electron becomes stronger and more energy is required to remove an electron from the anion, hence having higher ionization energy.

Successive ionization energies of an atom

It is possible to remove all electrons from a gaseous atom as long as sufficient energy is available.

Successive ionisation energies are the subsequent ionisation energies after the removal of the first electron from an atom (first ionisation energy).

Trend: The successive ionization energies increase as the electrons are removed (from the 1st ionization energy to the last ionization energy). i.e. Successive ionisation energies increase with increase in ionisation number.

Explanation: This is because as electrons are removed, the number of electrons reduces but the number of protons remains constant and the proton to electron ratio increases. This leads to increase in the nuclear attraction for the remaining electrons (i.e. the remaining electrons become more strongly attracted towards the nucleus). Therefore, the remaining electrons require more energy to be removed.

e.g. Beryllium atom with four electrons has four successive ionization energies.

$Be_{(g)} \rightarrow Be^{+}_{(g)} + e^{-}$ 1st ionization energy = +900 kJ/mol

$Be^{+}_{(g)} \rightarrow Be^{2+}_{(g)} + e^{-}$ 2nd ionization energy = +1758 kJ/mol

$Be^{2+}_{(g)} \rightarrow Be^{3+}_{(g)} + e^{-}$ 3rd ionization energy = +14905 kJ/mol

$Be^{3+}_{(g)} \rightarrow Be^{4+}_{(g)} + e^{-}$ 4th ionization energy = +21060 kJ/mol

Beryllium has thus 4 electrons, two of which are not easy to remove i.e. the 3rd and 4th electrons.

To remove the 3rd electron, there is a sharp or sudden (very big) increase in ionisation energy. Therefore, the 3rd electron is being removed from an inner and thermodynamically stable energy level (different energy level) of electrons. Thus the 1st two electrons are in the outermost energy level hence being in group II.

i.e. Beryllium has two electrons in the higher energy level which are easy to remove and two other electrons in the lower energy level which is stable and very near to the nucleus hence difficult to remove.

Note that:

(a) Alkali metals (group IA) have only one electron in the outermost shell. They easily lose the electron and hence form ions most readily. Therefore, they have the lowest 1st ionization energies and are consequently the most reactive metals.

However, noble gases (group O) have completely (fully) filled outermost energy levels. They are stable and cannot form ions easily. Therefore, they have the highest ionization energies.

(b) Transition or d-block elements have similar (almost the same) ionization energies, across the periods. This is because as the nuclear charge increases, it is balanced by the screening effect caused by addition of an electron in the inner energy levels. This is clearly shown in the electron distribution of electrons in atoms of elements of the d-block.

Variation of 1st ionization energy across periods 2 and 3 of the Periodic table

Period 2

Element	Li	Be	B	C	N	O	F	Ne
1st I.E ($kJmol^{-1}$)	+520	+899	+801	+1086	+1403	+1314	+1681	+2080
Atomic No.	3	4	5	6	7	8	9	10

Period 3

Element	Na	Mg	Al	Si	P	S	Cl	Ar
1st I.E ($kJmol^{-1}$)	+494	+737	+577	+786	+1012	+999	+1255	+1521
Atomic No.	11	12	13	14	15	16	17	18

Task: Plot separate graphs of first ionization energy against atomic number of Period 2 and Period 3 elements.

Trend: Generally, the first ionization energy increases across elements of Period 2 and 3. However, beryllium and nitrogen in period 2 and magnesium and phosphorus in period 3 have abnormally higher values.

Explanation: This is because on moving across the period from one element to the next, there is addition of an electron to the same outermost energy level which leads to a slight increase in the screening effect. There is also addition of a proton to the nucleus which leads to an increase in the nuclear charge. However, the increase in the nuclear charge is greater than the increase in screening effect hence the effective nuclear charge increases. Thus, the nuclear attraction for the outermost electrons becomes stronger so the amount of energy required to remove electrons from gaseous atoms of the elements increases hence increase in the first ionization energy.

Alternatively, on moving across the period from one element to the next, the atomic radius decreases. This decreases the distance between the nucleus and the outermost electrons. Thus, the nuclear attraction for the outermost electrons becomes stronger so that the amount of energy required to remove electrons from gaseous atoms of the elements hence increase in the first ionization energy.

However, beryllium ($1s^22s^2$) and nitrogen ($1s^22s^22p^3$) in Period 2 have exceptionally higher first ionization energies because of their stable electronic configurations. The first electron in beryllium is removed from a fully filled 2s-orbital while that in nitrogen is removed from half-filled 2p orbitals (sub energy level) both of which are thermodynamically (energetically) stable hence large amounts of energy are required to effect removal of an electron from these orbitals.

Similarly, the first ionization energies of magnesium ($1s^22s^22p^63s^2$) and phosphorus ($1s^22s^22p^63s^33p^3$) are exceptionally higher because of their thermodynamically stable electronic configurations. The first electron in magnesium is removed from a fully filled 3s-orbital while that in phosphorus is removed from half-filled 3p orbitals (sub energy level) both of which are thermodynamically (energetically) very stable hence large amounts of energy are required to effect removal of an electron from these orbitals.

Note that:

(a) The first ionization energy of boron ($1s^22s^22p^1$) is less than that of beryllium ($1s^22s^2$) in period 2 yet boron has a smaller atomic radius because the:-

(i) first electron in boron is removed from less stable 2p orbitals (sub energy level) because the orbitals are neither half-filled nor fully filled.

(ii) single electron in the 2p orbitals of the boron atom is more screened from the nuclear charge compared to the 2s orbital electrons in beryllium.

(iii) 2p orbital electron in boron has a lower penetrating power than the 2s orbital electrons in beryllium.

The combination of the above 3 reasons implies that less energy is required to remove the 2p orbital electron in boron than the energy required to remove the 2s orbital electron in beryllium.

Similar reasoning explains why aluminium ($1s^22s^22p^63s^23p^1$) has a lower 1st ionization energy than magnesium ($1s^22s^22p^63s^2$) in period 3.

(b) Oxygen ($1s^22s^22p^4$) has a lower first ionization energy than nitrogen ($1s^22s^22p^3$) in period 2. This is because the 2p orbitals (sub energy level) in the oxygen atom are neither half-filled nor fully filled. Therefore, it is less stable than the half-filled 2p orbitals of the nitrogen which contain 3 electrons. Hence less energy is required to remove the first electron from the gaseous atom of the oxygen.

Similar reasoning explains why sulphur ($1s^22s^22p^63s^23p^4$) has a lower 1st ionization energy than phosphorus ($1s^22s^22p^63s^23p^3$) in Period 3.

Sample Items:

1. At a secondary school laboratory in Wakiso District, a group of learners is investigating the properties of elements in Period 3 of the periodic table. During the experiment, they use computer simulations to study how atoms of magnesium and aluminium lose electrons to form ions.

They observe that removing the first electron from a magnesium atom requires more energy than removing the first electron from an aluminium atom. However, when they attempt to remove a second electron, they find that aluminium now requires more energy than magnesium. The learners are confused by this trend and turn to you for clarification.

Task: As a chemistry student, use your knowledge of atomic structure and electron configuration to explain to the learners:

(a) Why magnesium has a higher first ionization energy than aluminium.

(b) Why aluminium has a higher second ionization energy than magnesium.

In your explanation, relate your answer to electron arrangement, stability of orbitals, and the concept of shielding.

2. During a science exhibition at a school in Kampala, learners present a project comparing the ionization energies of elements in Period 2. One group focuses on nitrogen and oxygen and uses charts to illustrate how much energy is required to remove electrons from each atom.

They notice an unusual pattern: oxygen requires less energy than nitrogen to remove the first electron, yet more energy than nitrogen to remove the second electron. This observation seems to contradict the general trend across the period. The group asks you, as a chemistry student, to help them interpret their findings.

Task: As a chemistry student, explain to the group why oxygen has a:

(a) lower first ionization energy than nitrogen.

(b) higher second ionization energy than nitrogen.

In your response, include reference to electron pairing, orbital arrangement, and stability of half-filled subshells.

Variation of first ionization energy down a group of elements in the Periodic table

Group I

Element	Li	Na	K	Rb	Cs
First ionization energy (kJmol$^{-1)}$	520	494	419	403	376

Group II

Element	Be	Mg	Ca	Sr	Ba
First ionization energy (kJmol$^{-1)}$	899	737	590	549	503

Trend: The first ionization energy decreases down the group of elements in the Periodic Table.

Explanation: This is because on moving down the group from one element to the next, there is addition of a new **inner** energy level completely filled with electrons which leads to a large increase in the screening effect. There is also addition of protons to the nucleus which leads to an increase in the nuclear charge.

However, the increase in the screening effect is greater than the increase in nuclear charge hence the effective nuclear charge decreases. Thus, the nuclear attraction for the outermost electrons becomes weaker resulting into a decrease in the amount of energy required to remove an electron from a gaseous atom of the element hence decrease in the first ionization energy.

Alternatively, on moving down the group from one element to the next, the atomic radius increases. This increases the distance between the nucleus and the outermost electrons. Thus, the nuclear attraction for the outermost electrons becomes weaker so that the amount of energy required to remove electrons from gaseous atoms of the elements decreases and hence decrease in the first ionization energy.

Application of ionization energies in determining the chemistry of an element

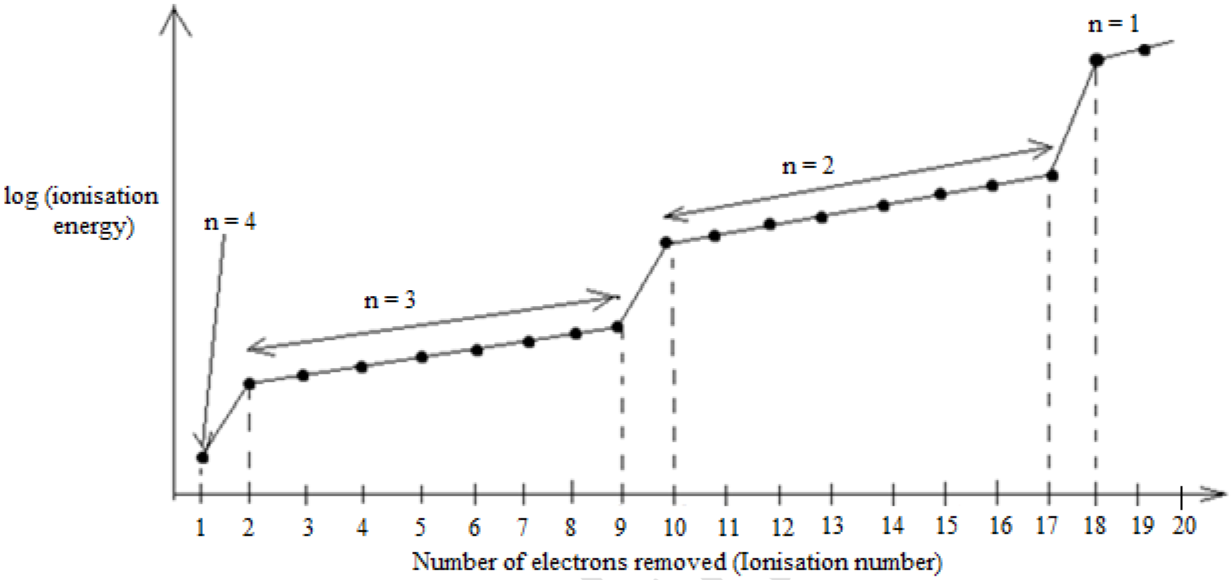
Ionization energies provide the following information about elements:

(1) **Atomic number of elements.**

The number of all successive ionization energies of an element indicates the number of electrons present in its atom hence the number of protons in its nucleus which is the atomic number of the element. e.g. Sodium has 11 successive ionization energies implying that the sodium atom has eleven electrons and eleven (11) protons hence sodium has an atomic number of 11.

(2) **Arrangement (distribution) of electrons in an atom.** This enables the writing of the electronic configuration of an atom.

This information is obtained effectively by plotting a graph of log (successive ionization energy) against number of electrons removed shows distinct breaks. **e.g.** The graph below is obtained for potassium.



The graph shows that each electron requires more energy to be removed than the previous one i.e. successive ionization energies increase with number of electrons removed in the order

1st I.E < 2nd I.E < 3rd I.E < 4th I.E etc

This is because of two reasons;

(a) When electrons are removed in succession (one after the other) from gaseous atoms or cations, the nuclear charge or number of protons in the nucleus of the cations formed remains constant but the number of electrons decreases therefore the effective nuclear charge (the ratio of number of protons to number of electrons of the cations formed) increases, therefore the nuclear attraction for the remaining electrons increases hence the amount of energy required to remove the next (subsequent) electron increases.

(b) The removal of the next electron may involve breaking into another energy level which is completely (fully) filled with electrons and therefore thermodynamically (energetically) stable and closer to the nucleus. Hence the electrons are more strongly attracted by the nucleus and so requiring a large amount of energy to be removed.

Note that:

(i) In addition to this trend of more energy required to be removed, there are some distinct breaks in the graph indicating that electrons are contained in different energy levels with the first energy level (n = 1) holding 2 electrons with fairly similar ionization energies both of which are very close to the nucleus hence very difficult to remove and have very high ionization energies.

The 2nd energy level (n = 2) has 8 electrons with fairly close ionization energies but distinctly less than the previous 2.

The 3rd energy level (n = 3) also has 8 electrons and the 4th energy level (n = 4) has 1 electron which is very easy to remove. i.e. requires the least amount of energy to be removed since it is in the outermost energy level.

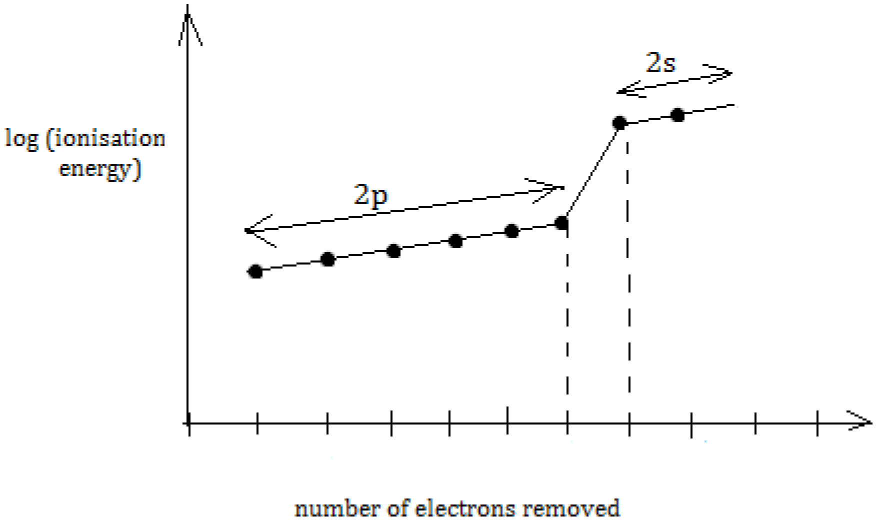
Therefore the 19 electrons in the potassium atom are arranged as follows:

Energy level	n = 1	n = 2	n = 3	n = 4
Number of electrons	2	8	8	1

Since the potassium atom shows 4 energy levels, it therefore belongs to period 4.

(ii) A **careful** plot of log (successive ionization energies) can also provide information about presence of sub-energy levels (sub shells).

e.g. A plot of log (successive ionization energies) in the 2nd energy level of potassium against number of electrons removed shows that there are 2 electrons with fairly the similar ionization energies which are nearer to the nucleus and they are followed by 6 electrons with also similar ionization energies but less than the first 2 electrons. i.e. **Graph of log (ionization energy) for the 2nd energy level for potassium atom**



From the above information, it indicates that the second energy level is further subdivided into 2 sub-energy levels; the 2s and the 2p.

Sub energy level	2s	2p
Number of electrons	2	6

Hence the 2nd energy level is further subdivided into 2 sub-energy levels i.e. $2s^2$ and $2p^6$.

Sample Item:

At a well-equipped government secondary school in Kampala, Senior Five learners are carrying out an inquiry-based chemistry investigation during their study of atomic structure and the periodic table. Their teacher has challenged them to identify an unknown element using experimental data obtained from a modern spectroscopic instrument at a nearby university laboratory.

The instrument measures the energy required to remove electrons one at a time from a gaseous atom of the element. The learners are reminded that this process is explained by the concept of Ionisation Energy, and that patterns in successive ionisation energies can reveal important information about the arrangement of electrons in an atom.

Working in groups, the learners carefully record the successive ionisation energies of the unknown element in the table below:

Ionisation number	**1st**	**2nd**	**3rd**	**4th**	**5th**	**6th**	**7th**	**8th**	**9th**	**10th**	**11th**
Ionisation energy (kJ/mol)	496	4560	6910	9543	13354	16613	20117	25496	28932	141362	159076

As they analyze the data, the teacher encourages them to apply their knowledge of Chemistry and to use graphical methods to better understand the trend in energy changes. The learners recall that **taking logarithms** can make large variations in data easier to interpret visually.

They are tasked with presenting their findings to the class and justifying their conclusions about the identity and position of the element in the periodic table.

Task: As a chemistry student,

(a) Using the data provided, plot a graph of $\log_{10}$ (ionisation energy) against ionisation number.

(b) By carefully analyzing the trend and any significant changes in the graph, determine the period and group in the Periodic Table to which the element belongs.

(c) Explain your reasoning based on the pattern of successive ionisation energies.

(3) **The number of energy levels and the Period to which an element belongs.** This can be got by the formula, n+1, where n is the number of sharp increases on the graph.

(4) **Determining the group in the Periodic Table to which an element belongs.**

This can be got by determining successive ionisation energies that correspond to the total number of electrons in the outermost energy level, before the first sharp increase in ionisation energy. i.e.

Element	1st I.E	2nd I.E	3rd I.E	4th I.E	5th I.E	6th I.E	7th I.E
Sodium	496	**4560**					
Magnesium	738	1450	**7730**				
Aluminium	578	1820	2750	**11600**			
Silicon	786	1580	3250	4360	**16100**		
Phosphorus	1012	1900	2910	4960	6270	**22200**	
Sulphur	1000	2250	3360	4560	7010	8500	**27100**
Chlorine	1251	2300	3820	5160	6540	9460	11000
Argon	1521	2670	3920	5770	7240	8780	12000

It is done by identifying the 1st set of two (a pair of) successive ionization energies having the largest energy difference and show the 1st sharp increase in ionisation energy as illustrated in the following **sample items**:

(a) At a secondary school laboratory in Uganda, a group of learners is carrying out an investigation on atomic structure and trends in the Periodic Table. Their teacher provides them with data for an unknown element Y, obtained from a scientific database. The data shows the energies required to remove electrons one after another from an isolated gaseous atom of element Y. The values recorded are as follows: the first ionization energy is 860 $kJmol^{-1}$, the second ionization energy is 1451 $kJmol^{-1}$, the third ionization energy is 7740 $kJmol^{-1}$, and the fourth ionization energy is 10,500 $kJmol^{-1}$.

The learners carefully analyze the pattern in the ionization energies and notice a significant increase between some of the successive values. They recall that such patterns can help determine the number of electrons in the outermost energy level of an atom.

Task: As a learner studying this topic:

(i) Identify the group of the Periodic Table to which element Y belongs.

(ii) Explain your answer based on the pattern observed in the ionization energy values.

Suggested Response:

(i) Y is in Group (II).

(ii) This is because there is a very big or sharp difference between the 2nd and 3rd ionization energies compared to differences between other pairs of successive ionization energies.

This implies that the 3rd electron was removed from an inner completely (fully) filled energy level which is thermodynamically stable. Thus the 1st two electrons are removed from the outermost energy level (valence) and the element belongs to group (II).

(b) At Midland Girls' School, a group of Senior Four learners is carrying out a research project in their chemistry laboratory on patterns in the Periodic Table. Their teacher provides them with experimental data for an unknown element X, including its first eight successive ionization energies (in kJ mol^{-1}) as follows: 786, 1580, 3230, 4360, 16000, 20000, 23000, and 29100.

The learners carefully study the data and notice a significant jump in ionization energy after the fourth electron is removed. They recall that such a sharp increase often indicates the removal of an electron from a more stable inner shell, which can help identify the position of an element in the periodic table.

As part of their discussion, they are also reminded of how elements form compounds, particularly chlorides, based on their valency and outer shell electron arrangement. They decide to use this information to predict how element X behaves chemically.

Task: As a chemistry student,

(i) Based on the ionization energy data provided, identify the group in the Periodic Table to which element X belongs. Give a reason for your answer.

(ii) Deduce the outer shell electronic configuration of element X.

(iii) Predict the formulae of two chlorides that element X can form.

Suggested Response:

(i) It is Group (IV), this is because there is a very big difference between the 4th and 5th ionization energies compared to other differences between successive ionization energies.

This implies that the 5th electron is removed from the next fully (completely) filled and thermodynamically stable energy level. Therefore, the first four electrons were removed from the outermost energy level and the element belongs to group (IV).

(ii) ns^2 np^2 where n is the principal quantum number.

(iii) XCl_2 and XCl_4.

(5) **Determining whether an element is either metal or non-metal.**

The magnitude of ionization energies is used as a measure of the metallic or non-metallic character of an element. Metals easily lose electrons while nonmetals easily gain electrons.

Generally, the 1st ionization energies of metals are close to and below 800 $kJmol^{-1}$ while those of non-metals are generally above 1000 $kJmol^{-1}$.

Explanation:

Down the group, there is increase in atomic radius such that there is a longer distance between the outermost electrons and the nucleus. Thus, ionization energies decrease such that the elements readily lose their valence electrons and become more metallic.

In group IV and V, there is change from metallic to non-metallic character.

Across the period, there is decrease in atomic radius as the outermost electron becomes closer to the nucleus. The ionisation energies increase as elements less readily lose the valence electrons or more readily gain electrons hence elements become less metallic or more nonmetallic.

Practice Items:

1. At a Kampala Senior School, a group of Senior Five learners is carrying out an advanced chemistry investigation on atomic structure and periodic trends in the laboratory. Their teacher has provided them with data for four unknown elements labeled A, B, C, and D. The students are tasked with analyzing the data to identify patterns in ionization energies and relate them to the position of elements in the periodic table.

The table below shows the **1st, 2nd, and 3rd ionization energies** of the elements:

Element	1st IE ($kJmol^{-1}$)	2nd IE ($kJmol^{-1}$)	3rd IE ($kJmol^{-1}$)
A	780	1500	7730
B	1315	3450	5300
C	580	1815	4620
D	500	4560	6900

The learners notice that some elements show a gradual increase in ionization energy, while others show a sudden large jump between successive values.

Task: As a chemistry student;

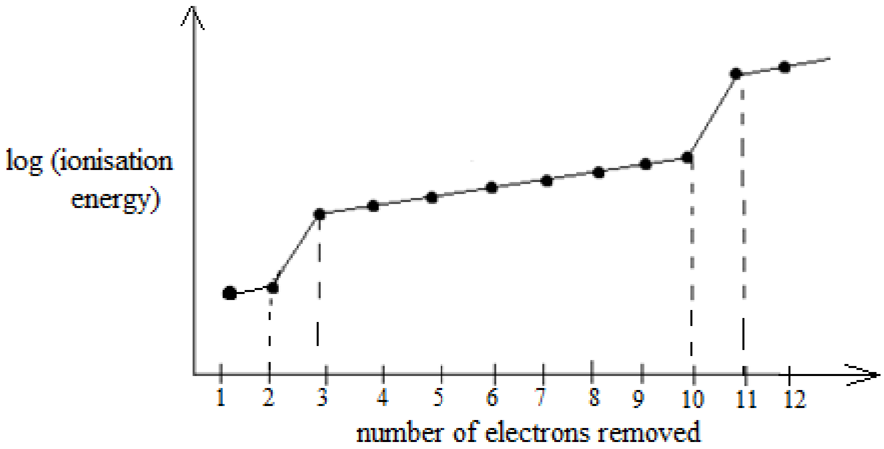
(a) Based on the information provided:

(i) Which one of the elements would you expect to exhibit **typical metallic characteristics**? Explain your answer.

(ii) Which of the elements would you classify as having **partly metallic (metalloid-like) characteristics**? Give reasons.

(b) Using the patterns observed in the ionization energies, help the learners determine the **likely groups in the periodic table** to which elements A, B, C, and D belong. Explain your reasoning clearly.

2. In an investigation, students at a school in Kawempe are given a graph showing the relationship between the **logarithm of successive ionization energies** and the **number of electrons removed** from an atom of another unknown element W.



During their discussion, one student recalls the principles of ionization energy and how it changes across periods and down groups.

Task: As a chemistry learner;

(a) With reference to the graph:

(i) Identify the **group** to which element W belongs. Give a reason based on the pattern observed.

(ii) State the **period** of element W and explain your answer.

(iii) Suggest the **identity of element W**.

(b) The graph shows a **sudden sharp increase** in energy required to remove the third electron. Explain why this occurs.

(c) Help the learners understand atomic changes by explaining:

(i) How the **size (atomic radius)** of element W changes as successive electrons are removed.

(ii) What would be the **sign of the energy change** if an electron is added to a neutral atom of W to form W^-.

(iii) How the addition of an electron (in part (ii)) would affect the **radius of the atom**.

(d) Based on your understanding of ion formation and stability, advise the learners whether element W can form **stable compounds containing W^+ ions**. Support your answer with a clear explanation.

3. ELECTRON AFFINITY

Electron affinity is the enthalpy change (heat change) that occurs when one mole of electrons is added to one mole of isolated / free gaseous atoms or anions of an element to form one mole of isolated gaseous anions.

Alternatively, electron affinity is the enthalpy change (heat change) that occurs when an electron is added to an isolated / free gaseous atom or anion of an element to form an isolated /free gaseous anion.

Types of electron affinity

(a) **First Electron Affinity.** This is the heat energy evolved when one mole of electrons is added to one mole of isolated gaseous atoms of an element to form one mole of singly charged (univalent) gaseous anions (uninegatively charged gaseous ion).

Alternatively, first electron affinity is heat energy released when an electron is added to an isolated gaseous atom to form a singly charged gaseous anion.

i.e. $M_{(g)} + e \rightarrow M^-_{(g)}$ $\Delta H = 1^{st}$ E.A ($kJmol^{-1}$)

e.g. For oxygen, $O_{(g)} + e \rightarrow O^-_{(g)}$ $\Delta H = -142\ kJmol^{-1}$

For sulphur, $S_{(g)} + e \rightarrow S^-_{(g)}$ $\Delta H = -200\ kJmol^{-1}$

N.B: (i) When tasked to define the term electron affinity, a candidate is advised to state the definition of first electron affinity.

(ii) For most elements, the 1st electron affinity is negative implying that the addition of an electron to a gaseous atom of the element is an ***exothermic process*** (accompanied by release of heat energy to the surroundings). This is so because the electron is being added to a *neutral gaseous atom*. Hence there is a *strong nuclear attraction for the electron being added* to the gaseous atom and heat energy is released.

(b) **Second Electron Affinity.** This is the heat energy absorbed when one mole of electrons is added to one mole of isolated singly charged (univalent) gaseous anions to form one mole of isolated doubly charged (divalent) gaseous anions.

i.e. $M^-_{(g)} + e \rightarrow M^{2-}_{(g)}$ $\Delta H = 2^{nd}$ E.A ($kJmol^{-1}$)

e.g. For oxygen, $O^-_{(g)} + e \rightarrow O^{2-}_{(g)}$ $\Delta H = +702\ kJmol^{-1}$

For sulphur, $S^-_{(g)} + e \rightarrow S^{2-}_{(g)}$ $\Delta H = +332\ kJmol^{-1}$

N.B: (i) The second electron affinity for any element is always positive implying that addition of an electron to a univalent gaseous anion of an element is an ***endothermic process***.

This is so because the *added (incoming) negatively charged electron is repelled by the negative charge on the gaseous anion*. Therefore, *energy is absorbed in order to overcome the repulsive forces* and effect addition of the electron.

The 3rd, 4th, 5th and other electron affinities can also be defined and are all positive for the same reason as for the second electron affinity.

(ii) The more negative (exothermic) the electron affinity, the more stable the anion formed.

(iii) The more positive (endothermic) the electron affinity, the less stable the ion formed.

(iv) Usually, formation of divalent ions is endothermic thus univalent ions are more stable than divalent ions.

Factors That Affect the First Electron Affinity

These are concerned with the measure of attraction between the nucleus and the incoming (added) electron.

(a) **Atomic Radius.**

Trend: The smaller the atomic radius, the higher the electron affinity.

Explanation: When the radius of the atom is smaller, the distance between the nucleus and the incoming electron is shorter and the nuclear attraction for the incoming electron is stronger. Thus, more energy is given out or released on addition of the electron leading to high electron affinity.

Similarly, the larger the atomic radius, the lower the electron affinity. This is because, when the radius of the atom is larger, then the nuclear attraction for the added electron is weaker. Thus, less energy is given out or released after addition of the electron, hence the electron affinity is lower.

e.g. $O_{(g)} + e \rightarrow O^-_{(g)}$ $\Delta H = -142\ kJmol^{-1}$

$F_{(g)} + e \rightarrow F^-_{(g)}$ $\Delta H = -333\ kJmol^{-1}$

(b) **Nuclear charge.**

Trend: The higher the nuclear charge, the higher the electron affinity.

Explanation: When the nuclear charge of the atom is higher, then the nuclear attraction for incoming electron is stronger. Thus, more energy is given out or released after addition of the electron, hence the electron affinity is higher.

Similarly, the lower the nuclear charge, the lower the electron affinity. This is because when the nuclear charge is lower, the attraction for the incoming electron is weaker. Thus, less energy is released leading to a lower electron affinity.

$Na_{(g)} + e \rightarrow Na^-_{(g)}$ $\Delta H = -21\ kJmol^{-1}$

$S_{(g)} + e \rightarrow S^-_{(g)}$ $\Delta H = -200\ kJmol^{-1}$

(c) **Screening Effect (Shielding Effect).**

Trend: The higher the screening effect, the lower the electron affinity.

Explanation: When the screening effect is higher, then the nuclear attraction for the added electron is weaker. Thus, less energy is released after addition of the electron leading to a lower electron affinity value.

Similarly, the lower the screening effect, the higher the electron affinity. This is because when the screening effect is lower, then the nuclear attraction for the added electron is stronger. Thus, more heat energy is given out after addition of the electron hence having higher electron affinity value.

$Cl_{(g)} + e \rightarrow Cl^-_{(g)}$ $\Delta H = -348\ kJmol^{-1}$

$Br_{(g)} + e \rightarrow Br^-_{(g)}$ $\Delta H = -324\ kJmol^{-1}$

(d) **Electronic Configuration of the Atom.**

Atoms (or ions) with fully filled or half-filled orbitals (sub energy levels) have lower electron affinity values than expected as less energy is evolved while some energy is absorbed (have positive values).

This is because fully filled or half-filled orbitals are thermodynamically (energetically) stable while those that are neither half-filled nor fully filled are thermodynamically unstable.

Thus, addition of an electron to orbitals that are half filled or fully filled does not easily take place and either less energy is liberated or energy is absorbed to effect addition of an electron hence having less negative or positive electron affinities.

$Mg_{(g)} + e \rightarrow Mg^-_{(g)}$ $\Delta H = +67\ kJmol^{-1}$ $1s^2 2s^2 2p^6 3s^2$

$P_{(g)} + e \rightarrow P^-_{(g)}$ $\Delta H = -60\ kJmol^{-1}$ $1s^2 2s^2 2p^6 3s^2 3p^3$

Element	Na	Mg	Al	Si	P	S	Cl	Ar
1st E.A ($kJmol^{-1}$)	-21	+67	-30	-135	-60	-200	-348	0

Task: Plot a graph of first electron affinity against atomic number of Period 3 elements.

(e) **Penetrating Power of orbitals.**

Trend: The higher the penetrating power of orbitals, the higher the electron affinity.

Explanation: This is because when the penetrating power of orbitals is higher, the outer orbitals come closer to the nucleus. Thus, the nuclear attraction for the electron being added is stronger and it becomes easier for the incoming electron to be added hence more energy is released.

(f) **Net charge on an ion within the same period.**

(i) Anions. (Generally, have positive electron affinities)

Trend: The higher the net charge on the anion, the more positive the electron affinity. Alternatively, the lower the net charge on the anion, the less positive the electron affinity.

Explanation: When the net charge on an anion is higher, then the proton to electron ratio is lower and there is also a stronger repulsion by the negatively charged anion for the incoming electron. The nuclear attraction for the incoming electron becomes weaker and more energy is absorbed to overcome the repulsive forces to effect addition of an electron to the anion, hence having a more positive electron affinity.

Alternatively, when the net charge on the anion is lower, then the proton to electron ratio is higher and there is also a weaker repulsion of the negatively charged anion for the incoming electron. The nuclear attraction for the incoming electron becomes stronger and less energy is absorbed to overcome the repulsive forces to effect addition of an electron to the anion, hence having a less positive electron affinity.

e.g. For Phosphorus, $P^{-}_{(g)} + e \rightarrow P^{2-}_{(g)}$ $\Delta H = +468\ kJmol^{-1}$

$P^{2-}_{(g)} + e \rightarrow P^{3-}_{(g)}$ $\Delta H = +886\ kJmol^{-1}$

(ii) Cations. (Generally, have negative electron affinities)

Trend: The higher the net charge on the cation, the more negative the electron affinity. Alternatively, the lower the net charge on the cation, the less negative the electron affinity.

Explanation: When the net charge on a cation is higher, then the proton to electron ratio is also higher. The nuclear attraction for the incoming electron becomes stronger and more energy is released on adding an electron to the cation, hence having a more negative electron affinity.

Alternatively, when the net charge on the cation is lower, then the proton to electron ratio is also lower. The nuclear attraction for the incoming electron becomes weaker and less energy is released on adding an electron to the cation, hence having a less negative electron affinity.

Variation of First Electron Affinity Across Period 2 and 3 Elements of the Periodic Table

e.g. **Period 2**

Element	Li	Be	B	C	N	O	F	Ne
1st E.A ($kJmol^{-1}$)	-57	+66	-15	-121	+31	-142	-333	+99
Atomic number	3	4	5	6	7	8	9	10

Period 3

Element	Na	Mg	Al	Si	P	S	Cl	Ar
1st E.A ($kJmol^{-1}$)	-21	+67	-30	-135	-60	-200	-348	0
Atomic number	11	12	13	14	15	16	17	18

Task: Plot separate graphs of first electron affinity against atomic numbers for elements of Periods 2 and period 3 of the Periodic Table.

Trend: Generally, the first electron affinity increases across the Period 2 and Period 3 elements.

Explanation: This is because on moving across the period from one element to the next, there is addition of an electron to the same energy level which leads to a slight increase in the screening effect. There is also addition of a proton to the nucleus which leads to an increase in the nuclear charge. However, the increase in the nuclear charge is greater than the increase in screening effect hence the effective nuclear charge increases.

As a result, the nuclear attraction for the incoming electron becomes stronger and so the amount of heat energy released after addition of the electron to gaseous atom of the element increases hence increase in the first electron affinity.

Alternatively, on moving across the period from one element to the next, the atomic radius decreases. This decreases the distance between the nucleus and the outermost electrons. Thus, the nuclear attraction for the incoming electrons becomes stronger so that the amount of energy released when electrons are added to gaseous atoms of the elements increases hence leading to a general increase in the first electron affinity.

However, Beryllium, Nitrogen and Neon in Period 2 have positive first electron affinities because of their electronic configurations.

Beryllium $1s^22s^2$ has a fully filled 2s orbital which is thermodynamically stable and so it cannot accommodate the electron being added. The added electron is repelled by the stable 2s orbital electrons and also has to occupy the 2p orbitals which are slightly higher in energy compared to the 2s orbital. Therefore, energy must be absorbed to effect addition of the electron into the not easily accessible higher vacant 2p orbitals.

Nitrogen $1s^22s^22p^3$ has half-filled 2p orbitals that are thermodynamically stable. The electron is added to the half-filled 2p orbitals hence the 3 electrons in this orbital tend to repel the added electron. Therefore, energy is absorbed in order to overcome the repulsion and effect addition of the electron.

Neon $1s^22s^22p^6$ has 2p orbitals that are fully filled and are thermodynamically stable hence it cannot accommodate the added electron. The electron added is repelled by the stable 2p orbital electrons and also has to occupy the 3s orbital which is considerably higher in energy than the 2p orbitals. Therefore, energy must be absorbed to effect addition of the electron into the not easily accessible higher vacant 3s orbital.

Note that:

- The first electron affinity of magnesium in period 3 is positive for reasons similar to those of Beryllium.
- The first electron affinity of phosphorous in period 3 is less negative than expected for reasons similar to those of Nitrogen.
- The first electron affinity of argon in period 3 is zero for reasons similar to those of neon.
- The first electron affinity of atoms that have orbitals that contain one electron less to become half-filled or fully-filled have high electron affinities than expected. This is because by gaining an electron, it makes the electronic configuration to become half-filled and thermodynamically stable. e.g. silicon, $1s^22s^22p^63s^23p^2$ which has 2 electrons in the 3p orbitals gains an electron more easily and more exothermically. This is because by gaining the electron, it makes the 3p orbitals in silicon to have 3 electrons thus becoming half-filled and becoming thermodynamically stable.

The same reason applies to the first electron affinity of carbon and halogens (fluorine, chlorine).

Variation of First Electron Affinity Down a Group of Elements in the Periodic Table

e.g **Group VII**

Element	**F**	**Cl**	**Br**	**I**
1st Electron Affinity (kJmol^{-1})	-333	-348	-324	-294

Trend: Generally, the first electron affinity decreases down the group of elements.

Explanation: This is because on moving down the group from one element to the next, there is addition of a new **inner** energy level completely filled with electrons which leads to a large increase in the screening effect. There is also addition of protons to the nucleus which leads to an increase in the nuclear charge. However, the increase in the screening effect is greater than the increase in nuclear charge hence the effective nuclear charge decreases. Thus, the nuclear attraction for the added electrons becomes weaker and there is a decrease in amount of energy given out after addition of the electron to gaseous atoms of the elements.

Alternatively, on moving down the group from one element to the next, the atomic radius increases. This increases the distance between the nucleus and the outermost electrons. Thus, the nuclear attraction for the

incoming electrons becomes weaker so that the amount of energy released when electrons are added to gaseous atoms of the elements decreases and hence leading to a general decrease in the first electron affinity.
However, fluorine has a lower 1st electron affinity value than expected. This is because fluorine has the smallest atomic radius which reduces the distance between the lone pairs of electrons and the incoming electron. The electron density on the fluorine atom is thus higher (greater or more crowded) than that on the other halogens. Hence, the added electron experiences repulsion from the lone pairs of electrons in fluorine. Some energy is absorbed to overcome the repulsion as the nuclear attraction for the incoming electron is weaker in fluorine. Thus, less energy is given out after addition of an electron to a gaseous atom.

4. ELECTRONEGATIVITY

Electronegativity is the measure of the tendency/ability of a covalently bonded atom in a molecule to attract bonding electrons towards itself.
Therefore, electronegativity is a property of a covalently bonded atom and not an isolated atom e.g. in the hydrogen chloride molecule (H–Cl), the chlorine atom is more electronegative than the hydrogen atom, hence the pair of electrons in the hydrogen to chlorine covalent bond is attracted more by the chlorine atom than hydrogen atom. Consequently, the chlorine atom acquires a partial negative charge (δ^-) because the electron density is displaced towards it while the hydrogen atom acquires a partial positive charge (δ^+) forming a dipole. Under these conditions, the covalent bond in between hydrogen and chlorine atoms is said to be polar.
A polar covalent bond is one in which electrons are not equally shared between the two bonded atoms due to difference in electronegativity of the atoms.
Polarization introduces ionic character in a covalent bond and the larger the difference in electronegativity of the bonded atoms, the more polar the covalent bond is and the greater the ionic character in the covalent bond.
However, covalent bonds formed between atoms of the same element are non-polar because the atoms have the same ability to attract electrons of the covalent bond.
Examples of molecules with non-polar covalent bonds include:
Chlorine Cl_2 (Cl – Cl), hydrogen H_2 (H – H), oxygen O_2 (O = O), nitrogen N_2 (N ≡ N) etc.
N.B: (i) Most non-metals are electronegative since they gain electrons to form anions.
The more electronegative a non-metallic atom is, the more chemically active it is.
(ii) Electronegativity involves both ionization and electron affinity since it is a measure of the power of an atom to attract electrons.
(iii) A pure covalent bond is formed when the atoms bonded have the same electronegativity.
e.g. In homonuclear diatomic molecules (molecules composed of one type of element).
(iv) Ionic character is developed when two atoms bonded have a large difference in their electronegativities.

Factors that Affect Electronegativity

These are concerned with the measure of attraction between the nucleus and the electrons in the covalent bond.

(a) **Atomic radius**

Trend: The smaller the atomic radius, the higher the electronegativity.
Explanation: When the radius of an atom in a covalent bond is smaller, the distance between its nucleus and electrons of the covalent bond is shorter. Thus, the nuclear attraction for the bonded electrons is stronger resulting into a higher electronegativity.
Similarly, the larger the atomic radius, the lower the electronegativity. This is because when the radius of a covalently bonded atom is larger, then the distance between the nucleus of the atom and the electrons of the

covalent bond is longer. Hence the nuclear attraction for the bonded electrons is weaker resulting into a lower electronegativity.

(b) **Nuclear charge**

Trend: The higher the nuclear charge, the higher the electronegativity.

Explanation: When the nuclear charge of an atom in a covalent bond is higher, there is stronger nuclear attraction for bonding electrons hence the higher the electronegativity.

Similarly, the lower the nuclear charge, the lower the electronegativity. This is because when the nuclear charge of an atom in a covalent bond is lower, there is a weaker nuclear attraction for bonding electrons hence the lower the electronegativity.

(c) **Shielding/Screening effect**

Trend: The higher the screening effect, the lower the electronegativity.

Explanation: When the screening effect of an atom in a covalent bond is higher, there is a weaker nuclear attraction for bonding electrons hence the lower the electronegativity.

Similarly, the lower the screening effect, the higher the electronegativity. This is because when the screening effect of an atom in a covalent bond is lower, there is a stronger nuclear attraction for bonding electrons hence the higher the electronegativity.

Variation of Electronegativity Across Elements of Period 2 and 3 of The Periodic Table

Period 2

Element	Li	Be	B	C	N	O	F
Electronegativity value	1.0	1.5	2.0	2.5	3.0	3.5	4.0

Period 3

Element	Na	Mg	Al	Si	P	S	Cl
Electronegativity value	0.9	1.2	1.5	1.8	2.1	2.5	3.0

Trend: Electronegativity increases across elements of periods 2 and 3 of the Periodic table.

Explanation: This is because on moving across the period from one element to the next, there is addition of an electron to the same energy level which leads to a slight increase in the screening effect. There is also addition of a proton to the nucleus which leads to an increase in the nuclear charge.

However, the increase in the nuclear charge is greater than the increase in screening effect hence the effective nuclear charge increases.

This results into an increase in the nuclear attraction for electrons in the covalent bond hence an increase in the electronegativity of the elements.

Alternatively, on moving across the period from one element to the next, the atomic radius decreases. This decreases the distance between the nucleus and the outermost electrons. Thus, the nuclear attraction for the electrons in the covalent bond increases hence leading to an increase in the electronegativity of the elements.

Variation of Electronegativity Down a Group of Elements in The Periodic Table

Group I

Element	Li	Na	K	Rb	Cs
Electronegativity value	1.0	0.9	0.8	0.8	0.7

Group VII

Element	F	Cl	Br	I
Electronegativity value	4.0	3.0	2.8	2.2

Trend: Electronegativity decreases down the group in the Periodic table.

Explanation: This is because on moving down the group from one element to the next, there is addition of a new **inner** energy level completely filled with electrons which leads to a large increase in the screening effect.
There is also addition of protons to the nucleus which leads to an increase in the nuclear charge.
However, the increase in the screening effect is greater than the increase in nuclear charge hence the effective nuclear charge decreases.
This results into a decrease in the nuclear attraction for electrons of the covalent bond hence a decrease in the electronegativity.
Alternatively, on moving down the group from one element to the next, the atomic radius increases. This increases the distance between the nucleus and the outermost electrons. Thus, the nuclear attraction for electrons of the covalent bond decreases hence leading to a decrease in the electronegativity.

5. ELECTROPOSITIVITY

Electropositivity is the measure of the tendency/ability of an atom of an element to lose its valence (outermost) electrons leading to formation of a cation.
The greater this tendency, the more electropositive the atom of the element is and the lower this tendency is, the less electropositive the atom of the element is.
Metal atoms readily lose electrons to form cations unlike non-metal atoms; therefore, metal atoms are more electropositive than non-metals.
Electropositivity can be described in covalent bonds as; the measure of the tendency for an atom of an element to lose bonding electrons.

Factors Which Affect Electropositivity

These are concerned with the measure of attraction between the nucleus and the electron(s) in the outermost energy level.

(a) **Atomic radius**

Trend: The smaller the atomic radius, the lower the electropositivity of the atom of the element.
Explanation: When the atomic radius is smaller, the distance between the nucleus and valence electrons is shorter. Therefore, the nuclear attraction for the valence electrons is stronger and so these electrons are not easily lost to form cations leading to lower electropositivity.
Similarly, the bigger the atomic radius, the higher the electropositivity of the atom of the element. This is because the distance between the nucleus and valence electrons is longer. Thus, the nuclear attraction for valence electrons is weaker and so electrons are readily lost to form cations leading to higher electropositivity.

(b) **Nuclear charge**

Trend: The higher the nuclear charge of an atom, the lower the electropositivity.
Explanation: When the nuclear charge is higher, the nuclear attraction of the outer most electrons is stronger. Hence the valence electrons are not readily lost to form cations leading to lower electropositivity.
Similarly, the lower the nuclear charge, the higher the electropositivity of the atom of an element. This is because nuclear attraction for the outermost electrons is weaker and so these electrons are readily lost to form cations leading to higher electropositivity.

(c) **Screening/Shielding effect**

Trend: The higher the screening effect in an atom, the higher electropositivity of the atom.
Explanation: When the screening effect is higher, the nuclear attraction for the outer most electron is weaker. Hence the valence electrons are readily lost to form positive ions leading to higher electropositivity.

Similarly, the lower the screening effect is an atom, the lower the electropositivity. This is because the nuclear attraction for the outermost electrons is stronger and so these electrons are not easily/readily lost to form cations leading to a lower electropositivity.

(d) **The electronic configuration of the atom**

Atoms with fully filled or half-filled orbitals have low electropositivity values.

This is because fully filled or half-filled orbitals are thermodynamically (energetically) stable while those that are neither half-filled nor fully filled are thermodynamically unstable.

Thus, loss of an electron from orbitals that are half filled or fully filled does not easily take place and the tendency to form cations is low hence having low electropositivity values.

Variation of Electropositivity Across a Period of Elements in The Periodic Table

Trend: Electropositivity decreases across elements of a period in the Periodic Table.

Explanation: This is because on moving across the period from one element to the next, there is addition of an electron to the same energy level which leads to a slight increase in the screening effect. There is also addition of a proton to the nucleus which leads to an increase in the nuclear charge. However, the increase in the nuclear charge is greater than the increase in screening effect hence the effective nuclear charge increases.

Therefore, the nuclear attraction for the outermost electrons increases and so tendency for these electrons to be lost to form cations decreases, hence decrease in electropositivity.

Alternatively, on moving across the period from one element to the next, the atomic radius decreases. This decreases the distance between the nucleus and the outermost electrons. Thus, the nuclear attraction for the outermost electrons increases and so tendency for these electrons to be lost to form cations decreases, hence decrease in electropositivity.

Variation of Electropositivity Down a Group of Elements in The Periodic Table

Trend: Electropositivity increases down a group of elements.

Explanation: This is because on moving down the group from one element to the next, there is addition of a new **inner** energy level completely filled with electrons which leads to a large increase in the screening effect.

There is also addition of protons to the nucleus which leads to an increase in the nuclear charge.

However, the increase in the screening effect is greater than the increase in nuclear charge hence the effective nuclear charge decreases.

This results into a decrease in the nuclear attraction for the outermost electrons. Therefore, the tendency for these electrons to be lost to form cations increases hence leading to increase in electropositivity.

Alternatively, on moving down the group from one element to the next, the atomic radius increases. This increases the distance between the nucleus and the outermost electrons. This results into a decrease in the nuclear attraction for the outermost electrons. Therefore, the tendency for these electrons to be lost to form cations increases hence leading to increase in electropositivity.

Other Physical Properties that vary in the Periodic Table include:

6. **MELTING POINTS** and **BOILING POINTS.**

Melting is the process by which the solid state of a substance is in equilibrium with its liquid state at a constant vapour pressure (so that their vapour pressures are equal).

Melting point of a substance is the constant temperature at which the liquid substance and the solid substance are in equilibrium so that their vapour pressures are equal.

Boiling is the process by which the saturated vapour pressure of a liquid substance is equal to the external atmospheric pressure at a constant temperature. (Compare with vapourisation).

Alternatively, boiling is the process by which the liquid state of a substance is in equilibrium with its vapour at a constant temperature.

Alternatively, boiling is the process by which a liquid substance is converted into its vapour at a constant temperature.

Boiling point of a substance is defined as the constant temperature at which the vapour pressure of the liquid is equal to the external atmospheric pressure.

Alternatively, boiling point is the constant temperature at which the liquid state of a substance is in equilibrium with its vapour.

Preamble: When a *solid chemical substance melts*, the forces of attraction between the particles are not completely broken but only *loosened or weakened enough* and the particles move around but not so freely. The forces of attraction between the particles are still present in the molten (liquid) state, and are not entirely broken until it boils. This means that boiling point but not melting point, are a much better guide to the real strengths of the solid substances.

When a *liquid chemical substance boils*, the forces of attraction between the particles are assumed to be *completely broken* and the particles move around so freely.

The type of particles involved are atoms, ions, molecules and electrons.

Melting points and **boiling points** involve the effect of heat energy on the forces of attraction between particles leading to a change in the physical state. i.e. physical change.

Melting point specifically is a measure of energy required to breakdown the regular arrangement of particles in a crystal lattice (solid structure). It varies according to the type of crystal formed by the element and it does not not uniformly vary in the Periodic Table.

Therefore melting and boiling points of chemical substances are governed by the *strength of the forces of attraction between the particles* in their structures.

A **crystal lattice** is the regular arrangement of atoms, ions or molecules in a crystal. The types of solid structures or crystal lattices in chemical substances include the:

- Giant ionic structure. This is shown by compounds formed between a metal and a non-metal.
- Giant molecular (atomic or covalent) structure. This is shown by a few nonmetals (such as carbon, silicon, germanium) and few covalent compounds (such as silicondioxide and germanium dioxide).
- Giant metallic structure. This is shown by metals.
- Simple molecular (atomic) structure. This is shown by most nonmetals and covalent compounds except those that form a giant covalent structure.

The types of forces of attraction between the particles in a structure include:

- Strong electrostatic forces as ionic bonds and metallic bonds shown by ionic compounds and metals respectively.
- Weak Van der Waal's forces between molecules shown by most nonmetals and covalent compounds.
- Strong hydrogen bonds between molecules shown by a few covalent compounds.
- Strong covalent bonds between atoms shown by a few nonmetals such as carbon and some covalent compounds such as silicondioxide.

7. METALLIC and NON-METALLIC CHARACTER.

The metallic character of an element depends on how easy an atom can lose electrons whereas the nonmetallic character of an element depends on how easy an atom can gain electrons.

Whether an element behaves like a metal or a non-metal can be determined using:

(a) Electropositivity.

The metallic or nonmetallic character of an element depends on how easy an atom can lose the valence electrons to form positively charged ions. Thus, the more electropositive an element is, the more metallic or less nonmetallic it is. Similarly, the less electropositive an element is, the less metallic or more non-metallic it is.

(b) Electronegativity.

The non-metallic or metallic character of an element depends on how easy an atom in a covalent bond can attract bonding electrons. Thus, the more electronegative an element is, the more non-metallic or less metallic it is. Similarly, the less electronegative an element is, the more metallic or less nonmetallic it is.

(c) Ionization energy.

The metallic or nonmetallic character of an element depends on how easy electrons can be removed from an atom.

The lower the ionisation energy the more metallic or less nonmetallic an element is. Metallic character is shown by elements with low values of ionisation energy. e.g. the first ionization energy of all metals is close to or below 800 kJ/mol.

Similarly, the higher the ionisation energy the more nonmetallic or less metallic an element is. Non-metallic character is shown by elements with high values of ionisation energy. e.g. the first ionization energy of all non-metals is way above 800 kJ/mol.

(d) Electron affinity.

The non-metallic or metallic character of an element depends on how easy an atom can gain electrons. Thus, the higher the electron affinity of an element, the more non-metallic or less metallic it is. Similarly, the lower the electron affinity of an element, the more metallic or less nonmetallic it is.

Non-metallic character is shown by elements with high values of electron affinity while metallic character is shown by elements with low values of electron affinity.

General Variation of Metallic and Non-metallic Character Across the Period:

Trend: There is a gradual change from metallic to non-metallic character across the Period.

i.e. Metallic character decreases while non-metallic character increases across elements of a period in the Periodic Table.

General Variation of Metallic and Non-metallic Character Down the Group:

Trend: The metallic character increases while non-metallic character decreases down the group in the Periodic Table.

8. ATOMISATION ENERGY.

This is the heat energy required (heat absorbed) when one mole of gaseous atoms is formed from its elements at standard conditions (in their normal physical states).

9. STANDARD ELECTRODE POTENTIAL (S.E.P).

This is the value of the reduction potential of an electrode (half-cell) measured using a standard hydrogen electrode under standard conditions (at 298K and 1 atmosphere when a metal electrode is dipped in a solution of ions at unit (1 molar) concentration).

Alternatively, this is the electromotive force (emf) of a cell in which the electrode (half-cell) is on the right-hand side while the standard hydrogen electrode is on the left-hand side under standard conditions.

10. CONDUCTANCE OF ELECTRICITY (ELECTRICAL CONDUCTANCE).

This is a property of a conductor which determines how easily electric current can go through it. Electrical conductance is generally high for metals because they have delocalised electrons and very low or negligible for non-metals except graphite.

11. HYDRATION ENERGY and HYDRATED RADIUS.

Hydration energy is the energy liberated when one mole of gaseous ions is completely surrounded by water molecules to form an infinitely dilute solution at standard conditions.
Alternatively, it is the heat energy evolved when one mole of gaseous ions dissolves in water to form an infinitely dilute solution at standard conditions.
Hydrated radius is half the inter nuclear distance between two identical ions in contact with each other in an infinitely dilute aqueous solution.

12. ANOMALOUS BEHAVIOUR AND DIAGONAL RELATIONSHIP.

Anomalous behaviour is the tendency of certain elements in Period 2 of the periodic table to show properties or characteristics which are different from other members of the same group. e.g; Li in group I, Be in group II, B in group III e.t.c.
Diagonal relationship is the tendency of certain elements in Period 2 to show properties which are similar to other elements which are immediately diagonally below them in period 3 of the adjacent group across the period.
It mainly exists between the following pairs of elements, Lithium and Magnesium, Beryllium and Aluminium, Boron and Silicon.
Moving one step horizontally increases the non-metallic character and moving a step vertically downwards increases metallic character. Therefore, moving diagonally, these tendencies compensate and cancel out each other resulting into similar properties.
e.g. The charge to radius ratio of the beryllium ion is similar to that of aluminium ion a diagonal neighbour leading to a similar chemistry.
Elements show ***anomalous behaviour*** because they have;

- the smallest atomic radii.
- the ability to form cations with the highest charge densities and polarising power.
  Thus form compounds with;
- high covalent character.
- high lattice energies and hence low solubilities.
- the highest electronegativities or lowest electropositivities.
- the highest positive (usually nonmetals) or highest negative (usually metals) standard electrode potentials.
- no d-orbitals.
- the ability to form anions with lowest polarisability.

Elements show ***diagonal relationship*** because they have similar:

- small atomic radii.
- small cations with similar high charge densities and similar high polarising power leading to formation of compounds with high degree of covalency.
- high electronegativity or low electropositivity.
- high negative or positive standard electrode potentials values (abnormally high negative or positive standard electrode potentials).

Diagonal Relationship Between Lithium and Magnesium

The two elements resemble each other as shown by the following properties.

(a) Their carbonates are unstable to heat and thus decompose to give oxides and carbondioxide.

$Li_2CO_{3(s)} \rightarrow Li_2O_{(s)} + CO_{2(g)}$ and $MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$

Note that; other group I metal carbonates are stable to heat and cannot decompose. Explain.

(b) They both react with carbon and nitrogen when heated to form dicarbides and nitrides respectively. i.e. $6Li_{(s)} + N_{2(g)} \rightarrow 2Li_3N_{(s)}$ and $3Mg_{(s)} + N_{2(g)} \rightarrow 2Mg_3N_{2(s)}$

Also; $2Li_{(s)} + 2C_{(s)} \rightarrow Li_2C_{2(s)}$ (Lithium dicarbide) and $Mg_{(s)} + 2C_{(s)} \rightarrow MgC_{2(s)}$

(c) Their dicarbides react with water giving off ethyne gas and their metal hydroxides.

i.e. $Li_2C_{2(s)} + 2H_2O_{(l)} \rightarrow 2Li(OH)_{(aq)} + HC\equiv CH_{(g)}$ or $Li_2C_{2(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + C_2H_{2(g)}$

and $MgC_{2(s)} + 2H_2O_{(l)} \rightarrow Mg(OH)_{2(s)} + C_2H_{2(g)}$

(d) Their chlorides posses covalent character. i.e. LiCl or $MgCl_2$. This is due to polarisation of the chloride ion in the ionic bond forming a partial covalent character.

Hence lithium chloride and magnesium chloride can dissolve in ethanol (an organic solvent) and both chlorides have very low melting points.

(e) They both form monoxides when heated in air. i.e;

$4Li_{(s)} + O_{2(g)} \rightarrow 2Li_2O_{s)}$ and $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

(Note that lithium does not form peroxides unlike other group I metals like sodium)

(f) Their nitrates readily decompose when heated to form normal oxides, nitrogen dioxide and oxygen.

$4LiNO_{3(s)} \rightarrow 2Li_2O_{(s)} + 4NO_{2(g)} + O_{2(g)}$ and $4Mg(NO_3)_{2(s)} \rightarrow 2MgO_{(s)} + 4NO_{2(g)} + O_{2(g)}$

(Note that other group I nitrates decompose to form metal nitrites and oxygen gas)

(g) Their ions become heavily hydrated in aqueous solutions e.g. $[Li(H_2O)_4]^+$, $[Mg(H_2O)_6]^{2+}$ and form acidic solutions. This is due to polarisation of the water molecules leading to hydrolysis and releasing hydrogen ions in the solution.

(h) Their carbonates and phosphates are insoluble in water. This is due to polarisation of the carbonate and phosphate ions forming a largely covalent character.

Question:

Jane is a high school student helping her cousin with a science project on alkali metals. As part of the experiment, they observe how lithium, sodium, and potassium react with water, air, and chlorine. Jane notices that lithium behaves differently compared to the others.

Task 1: *Based on Jane's observations, write* ***balanced chemical equations*** *for* ***four reactions*** *where lithium behaves differently from sodium and potassium. Explain the differences in each case.*

Task 2: *Using scientific reasoning,* ***explain why lithium's reactions differ*** *from those of the other alkali metals in the same group. Consider factors such as atomic radius, ionization energy, and charge density in your explanation.*

Diagonal Relationship Between Beryllium and Aluminium

The two elements resemble each other as shown by the following properties.

(a) Both metals are amphoteric hence react with both hot dilute non-oxidising acids (i.e. except nitric acid) and strong alkalis (e.g. sodium hydroxide) to liberate hydrogen gas i.e;

(i) acids: $2Al_{(s)} + 6H^+_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 3H_{2(g)}$ and $Be_{(s)} + 2H^+_{(aq)} \longrightarrow Be^{2+}_{(aq)} + H_{2(g)}$

(ii) alkalis: $2Al_{(s)} + 2OH^-_{(aq)} + 6H_2O_{(l)} \longrightarrow 2[Al(OH)_4]^-_{(aq)} + 3H_{2(g)}$

or $2Al_{(s)} + 2OH^-_{(aq)} + 2H_2O_{(l)} \longrightarrow 2AlO_2^-{}_{(aq)} + 3H_{2(g)}$

$Be_{(s)} + 2OH^-_{(aq)} + 2H_2O_{(l)} \longrightarrow [Be(OH)_4]^{2-}_{(aq)} + H_{2(g)}$

or $Be_{(s)} + 2OH^-_{(aq)} \longrightarrow BeO_2^{2-}{}_{(aq)} + H_{2(g)}$

(b) Their oxides and hydroxides are amphoteric. Thus for oxides they react with;

(i) acids as;

- oxides $Al_2O_{3(s)} + 6H^+_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 3H_2O_{(l)}$

  $BeO_{(s)} + 2H^+_{(aq)} \longrightarrow Be^{2+}_{(aq)} + H_2O_{(l)}$
- hydroxides $Al(OH)_{3(s)} + 3H^+_{(aq)} \longrightarrow Al^{3+}_{(aq)} + 3H_2O_{(l)}$

  $Be(OH)_{2(s)} + 2H^+_{(aq)} \longrightarrow Be^{2+}_{(aq)} + 2H_2O_{(l)}$

(ii) alkalis as;

- oxides $Al_2O_{3(s)} + 2OH^-_{(aq)} + 3H_2O_{(l)} \longrightarrow 2[Al(OH)_4]^-_{(aq)}$

  or $Al_2O_{3(s)} + 2OH^-_{(aq)} \longrightarrow 2AlO_2^-{}_{(aq)} + H_2O_{(l)}$

  $BeO_{(s)} + 2OH^-_{(aq)} + H_2O_{(l)} \longrightarrow [Be(OH)_4]^{2-}_{(aq)}$

  or $BeO_{(s)} + 2OH^-_{(aq)} \longrightarrow BeO_2^{2-}{}_{(aq)} + H_2O_{(l)}$
- hydroxides $Al(OH)_{3(s)} + OH^-_{(aq)} \longrightarrow [Al(OH)_4]^-_{(aq)}$

  or $Al(OH)_{3(s)} + OH^-_{(aq)} \longrightarrow AlO_2^-{}_{(aq)} + 2H_2O_{(l)}$

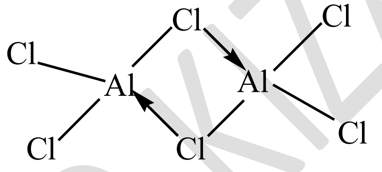
  $Be(OH)_{2(s)} + 2OH^-_{(aq)} \longrightarrow [Be(OH)_4]^{2-}_{(aq)}$

  or $Be(OH)_{2(s)} + 2OH^-_{(aq)} \longrightarrow BeO_2^{2-}{}_{(aq)} + 2H_2O_{(l)}$

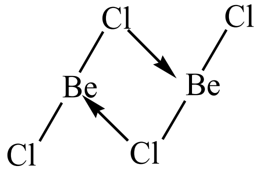
(c) Their chlorides and bromides are volatile covalent polymeric solids. This is due to polarisation of the chloride ion in the ionic bond forming a partial covalent character. Thus they are readily soluble (dissolves) in organic solvents such as ethanol to form dimers.

The chlorides also dimerise when the solids sublime.

e.g;



and



(d) Their soluble salts in aqueous solutions (e.g. chlorides) undergo hydrolysis releasing hydrogen ions (protons) that make the solutions to become acidic solutions.

Thus; $Al^{3+}_{(aq)} + 6H_2O_{(l)} \rightleftharpoons [Al(H_2O)_6]^{3+}_{(aq)}$

then hydrolysis; $[Al(H_2O)_6]^{3+}_{(aq)} + 3H_2O_{(l)} \rightleftharpoons Al(OH)_3(H_2O)_{3(aq)} + 3H_3O^+_{(aq)}$

Also; $Be^{2+}_{(aq)} + 4H_2O_{(l)} \rightleftharpoons [Be(H_2O)_4]^{2+}_{(aq)}$

then hydrolysis; $[Be(H_2O)_4]^{2+}_{(aq)} + 2H_2O_{(l)} \rightleftharpoons Be(OH)_2(H_2O)_{2(aq)} + 2H_3O^+_{(aq)}$

Dissolution of salts above results in formation of white fumes due to formation of aqueous hydrochloric acid which liberates hydrogen chloride gas.

Explanation: The beryllium ions in aqueous solutions of soluble salts exist as the tetraaquaberyllium ions, $[Be(H_2O)_4]^{2+}$. These ions are then hydrolysed to form acidic solutions.

Since the beryllium ion has a high charge density and thus a high polarising power. The beryllium to oxygen bond is strengthened and this weakens the oxygen to hydrogen bond of the water molecules. Consequently, the oxygen to hydrogen bond is broken releasing hydrogen ions in the solution leading to acidity. Other group II metal salts are not hydrolysed.

(e) Their carbides react with water and dilute mineral acids liberating methane gas. Thus;

$Be_2C_{(s)} + 3H_2O_{(l)} \longrightarrow 2Be(OH)_{2(s)} + CH_{4(g)}$ or $Be_2C_{(s)} + 4HCl_{(aq)} \longrightarrow 2BeCl_{2(aq)} + CH_{4(g)}$

and $Al_4C_{3(s)} + 12H_2O_{(l)} \longrightarrow 4Al(OH)_{3(s)} + 3CH_{4(g)}$ or $Al_4C_{3(s)} + 12HCl_{(aq)} \longrightarrow 4AlCl_{3(aq)} + 3CH_{4(g)}$

(f) Both metals are rendered passive to concentrated nitric acid due to formation of an inert and impervious film of oxides on the surface.

Note that; the examples used as characteristics or properties shown by diagonal relationhip are the same properties used for anomalous behaviour. e.g.

(a) Beryllium metal is amphoteric while other elements in group(II) are basic.

(b) Beryllium oxide and beryllium hydroxide are amphoteric while the oxides and hydroxides of other group(II) elements are basic.

(c) Beryllium chloride is a volatile covalent polymeric solid while other group(II) chlorides are non-volatile ionic solids.

(d) The soluble salts of beryllium e.g. beryllium chloride; are hydrolysed by water releasing hydrogen ions that lead to acidic solutions which is not the case for other group(II) salts.

(e) Beryllium carbide reacts with water and dilute mineral acids liberating methane gas while other group(II) metal carbides liberate ethyne gas. Magnesium forms another carbide that liberates propyne gas. Thus;

$Be_2C_{(s)} + 3H_2O_{(l)} \longrightarrow 2Be(OH)_{2(s)} + CH_{4(g)}$ and $Be_2C_{(s)} + 4HCl_{(aq)} \longrightarrow 2BeCl_{2(aq)} + CH_{4(g)}$

$MC_{2(s)} + 2H_2O_{(l)} \longrightarrow M(OH)_{2(s)} + C_2H_{2(g)}$ and $MC_{2(s)} + 2HCl_{(aq)} \longrightarrow MCl_{2(aq)} + C_2H_{2(g)}$

where M represents Mg, Ca, Sr and Ba.

Also for magnesium;

$Mg_2C_{3(s)} + 4H_2O_{(l)} \longrightarrow 2Mg(OH)_{2(s)} + C_3H_{4(g)}$ and $Mg_2C_{3(s)} + 4HCl_{(aq)} \longrightarrow 2MgCl_{2(aq)} + C_3H_{4(g)}$

Diagonal Relationship between Boron and Silicon

The two elements resemble each other as shown by the following properties.

(a) Both boron and silicon are typical non-metals, having high melting and boiling points.

(b) Both elements do not form cations and form only covalent compounds with other nonmetals.

(c) Both exist in amorphous and crystalline state and exhibit allotropy.

(d) Their chlorides are liquids that produce light white fumes in most air and are readily hydrolysed by water.

$BCl_{3\,(l)} + 3H_2O_{\,(l)} \rightarrow B(OH)_{3\,(aq)} + 3HCl_{(aq)}$ $SiCl_{4\,(l)} + 4H_2O_{\,(l)} \rightarrow Si(OH)_{4\,(aq)} + 4HCl_{(aq)}$

(e) Both form weak acids like boric acid, H_3BO_3 and silicic acid, H_2SiO_3.

(f) Both form binary compounds with several metals to give borides and silicides. These borides and silicides react with phosphoric acid to give mixture of boranes and silanes.

$3Mg_{(s)} + 2B_{(s)} \rightarrow \underset{\text{(Magnesium boride)}}{Mg_3B_{2\,(s)}}$ $Mg_3B_{2\,(s)} + H_3PO_{4\,(aq)} \rightarrow$ Mixture of boranes

$2Mg_{\,(s)} + Si_{\,(s)} \rightarrow \underset{\text{(Magnesium silicide)}}{Mg_2Si_{\,(s)}}$ $Mg_2Si_{\,(s)} + H_3PO_{4\,(aq)} \rightarrow$ Mixture of silanes

(g) Oxides of both elements are acidic and can be reduced by limited amount of magnesium. In excess of magnesium, the boride and silicide are formed.

$B_2O_{3\,(s)} + 3Mg_{\,(s)} \rightarrow 3MgO_{\,(s)} + 2B_{(s)}$ $SiO_{2\,(s)} + 2Mg_{\,(s)} \rightarrow 2MgO_{\,(s)} + Si_{\,(s)}$

(h) Both the elements and their oxides are acidic hence readily soluble in alkalis.

$2B_{\,(s)} + 6NaOH_{\,(aq)} \rightarrow 2Na_3BO_{3\,(aq)} + 3H_{2\,(g)}$

$Si_{(s)} + 2NaOH_{(aq)} + H_2O_{(l)} \rightarrow Na_2SiO_{3\,(aq)} + 2H_{2\,(l)}$

$B_2O_{3\,(s)} + 6NaOH_{(aq)} \rightarrow 2Na_3BO_{3\,(aq)} + 3H_2O_{(l)}$

$SiO_{2\,(s)} + 2NaOH_{(aq)} \rightarrow Na_2SiO_{3\,(aq)} + H_2O_{(l)}$

(i) Acids of both these elements form volatile esters on heating with an alcohol in presence of concentrated sulphuric acid.

$B(OH)_{3\,(aq)} + 3ROH_{(aq)} \rightarrow B(OR)_{3\,(aq)} + 3H_2O_{(l)}$

$Si(OH)_{4\,(aq)} + 4ROH_{(aq)} \rightarrow Si(OR)_{4\,(aq)} + 4H_2O_{(l)}$

(j) Both form several volatile and spontaneously catch fire on exposure to air (are inflammable). They are commonly called boranes and silanes. The hydrides are hydrolysed readily. Lower hydrides are obtained by reduction of chlorides with lithium aluminium tetrahydride, $LiAlH_4$.

$4BCl_{3\,(l)} + 3LiAlH_{4\,(l)} \rightarrow 3AlCl_{3\,(s)} + 3LiCl_{(s)} + 2B_2H_{6\,(g)}$

$SiCl_{4\,(l)} + LiAlH_{4\,(l)} \rightarrow AlCl_{3\,(s)} + LiCl_{(s)} + SiH_{4\,(g)}$

Practice Items:

1. (a) The science club is conducting a research project in your school on how different elements behave in various chemical reactions.

To present their findings, they are asked to create a **digital infographic or video presentation** explaining key atomic properties that influence chemical behavior.

Task: As a chemistry learner, help the science club to explain:

(i) What is meant by the terms; **ionization energy, electron affinity** and **atomic radius**. Highlight how atomic radius as a property is related to ionisation energy and electron affinity and how it affects the way elements interact or react.

(ii) Why removing each additional electron successively from an atom requires more energy.

(b) A materials analyst at a research institute is focused on advanced materials and their electronic properties.

The supervisor has asked analyst to analyze data for elements in the third period of the periodic table.

The focus is on understanding how electron affinity and melting point trends relate to atomic structure and bonding.

The table below summarizes data for elements sodium (Na) to argon (Ar).

Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
Atomic Number	11	12	13	14	15	16	17	18
1st Electron Affinity (kJ/mol)	-21	+67	-30	-135	-60	-200	-348	0
Melting Point (°C)	98	650	660	1429	44	120	-101	-189

Task: As a chemistry student, help the materials analyst to:

(i) Present a visual representation of how the first electron affinity changes across the third period by plotting a suitable graph.

(ii) Explain the shape of the graph as the graph plotted reveals a non-linear trend in electron affinities across the period.

(iii) Compare and explain the differences in melting points of sodium (Na), magnesium (Mg), silicon (Si), and phosphorus (P) since the team is exploring materials for heat-resistant applications and wants your input on why some elements melt at much higher temperatures than others.

In your response, refer to, the type of structure and bonding in each element and how these affect the melting point.

(c) A local science magazine is preparing an article titled “The Strange Behaviors of the Periodic Table.” You are invited to contribute a section explaining some of the unusual trends seen among elements.

Task: As a chemistry student, write a **short article** to explain what is meant by the terms; **anomalous behavior** and **diagonal relationship** in the periodic table.

Use a specific example with at least 3 reasons involving a group(II) element to illustrate your article.

2. (a) A laboratory assistant works at a local university. A group of new chemistry students is about to begin a unit on periodic trends, and the supervisor has asked the laboratory assistant to prepare a short guide to help them understand key terms used in atomic structure and periodicity.

Task: As a chemistry student, help the lab assistant to explain the meaning of the following scientific terms that he would include in the guide:

(i) Covalent radius.
(ii) First ionization energy.
(iii) Second electron affinity.
(iv) Electropositivity.

(b) A research team is developing a battery using alkali metals (Group I elements). You are part of the team, and your role is to advise on which metal is most suitable based on its periodic properties.

Task: As a chemistry student, use your understanding of trends in Group I metals to:

(i) Explain how **ionisation energy** changes as you move down the group and how this affects the reactivity of the metals.
(ii) Describe the trend in **electropositivity** and how it influences metal reactivity.
(iii) Discuss how **electrode potential** varies down the group and what it means for the battery's performance.
(iv) Examine the change in **melting point** down the group and how it could affect the battery design.

(c) You are tutoring a fellow student who is struggling to understand why certain elements in the periodic table do not follow expected trends in ionisation energy and chemical behaviour.

Task: As a chemistry student, help your friend by explaining the following observations:

(i) Why is the **first ionization energy of aluminium** lower than that of magnesium, despite aluminium having more protons?
(ii) Why is the **third ionization energy of magnesium** much higher than the second?
(iii) Beryllium, though a Group II element, behaves differently compared to others in its group. Explain this **anomalous behaviour**.

(d) During a chemistry debate club at school, one team argues that aluminium and beryllium, though in different groups, show similar chemical properties. The opposing team challenges this view.

Task: As a chemistry learner, to support your team you are required to:

(i) Define the term **diagonal relationship** in the periodic table.
(ii) Explain why elements such as beryllium and aluminium exhibit this type of relationship.
(iii) Using relevant **chemical equations**, demonstrate at least four ways in which aluminium's chemical behaviour is similar to that of beryllium.

3. (a) A mobile phone manufacturer is researching elements for creating efficient battery materials. While evaluating elements in period 2 of the periodic table, they observe that **beryllium seems to hold onto its outer electrons more tightly than boron**, despite boron coming after beryllium in the periodic table. This seems to go against the general trend of increasing ionization energy across a period.

Task: As a student researcher in chemistry, **explain why boron has a lower first ionization energy than beryllium**, even though ionization energy generally increases across a period.

(b) In a school lab, learners are doing a flame test using magnesium metal. When magnesium is heated, it emits a bright white flame. The teacher explains this happens when electrons absorb energy and jump to higher levels before falling back and releasing energy. A student wonders why it is more difficult to remove a second electron from magnesium than the first, even though both are from the same shell.

Task: Help the student to **explain why the second ionization energy of magnesium is higher than the first**, even though both electrons are in the same energy level.

(c) A water treatment engineer is using calcium salts to soften hard water. She explains that calcium typically forms Ca^{2+} ions in compounds. A student asks why calcium doesn't just form Ca^{+} ions, especially since removing only one electron requires less energy than removing two.

Task: Using your knowledge of ionization energy and atomic structure, **explain why calcium forms Ca^{2+} ions instead of Ca^{+} ions**, even though the second ionization energy is higher than the first.

(d) An air purifier manufacturer uses oxygen ions to remove pollutants. They discover that adding one electron to an oxygen atom releases energy, but adding a second electron requires energy input.

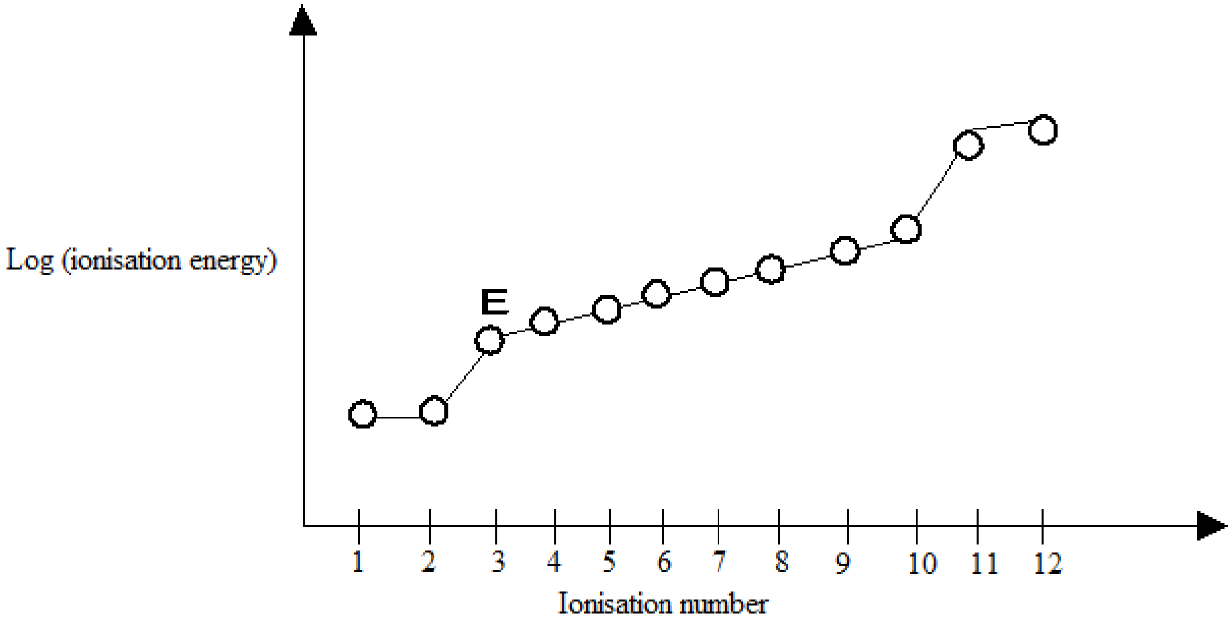
Task: As a chemistry consultant, **explain why the first electron affinity of oxygen is exothermic while the second is endothermic**, even though both involve gaining electrons.

4. (a) A high school student, Amina, is conducting research on the properties of elements for her chemistry project. She is particularly interested in how atoms lose electrons and how this affects their energy and size. During her research, she comes across data showing successive ionisation energies of an unknown element X, which shows a sharp increase at a certain point after several electrons have been removed. She also learns about concepts like ionisation energy and electron affinity, which are key to understanding chemical reactivity.

Amina comes across the terms; **ionisation energy** and **electron affinity** in her chemistry textbook.

Task: As a chemistry student, help Amina by defining the following terms; ionisation energy and electron affinity.

(b) The diagram below shows the successive ionisation energies of element X (not shown here), and Amina notices a sudden large jump in the energy required to remove the next electron after the second one.



Task: As a chemistry student, based on the pattern in ionisation energies:

(i) Which **group** in the Periodic Table does element X most likely belong to? Give a reason.

(ii) Which **period** is element X likely found in? Justify your answer. Suggest the **identity** of element X.

(iii) Explain to Amina why there is a **sudden increase** in the ionisation energy required to remove electron E.

(iv) Amina is curious about atomic size. Explain to her how the **atomic radius** of element X changes as electrons are successively removed.

(v) Amina wants to compare ionisation energies across a group. Help her understand how **ionisation energy changes down Group II** of the Periodic Table, and explain why this trend occurs.