2026

1^{ST} YEAR

CHEMISTRY

1st Year Chemistry

Chapter 1
Periodic Table and
periodic Properties

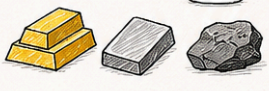
Muhammad Akhtar Jamil
MS Chemistry

One of the most important turning points in the history of science was the creation of the periodic table, which led to many important innovations. It is accurate to refer to the periodic table of elements as the “Symbol of Chemistry.” In the modern periodic table, 118 elements are arranged in tabular form in the current periodic table based on their atomic number.

visit website: www.everexams.com | subscribe youtube channel: chemscholar4u | Akhtar Jamil, MS Chemistry, 20 Years Experience

1.1 Historical Background

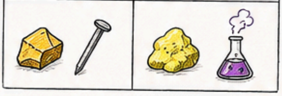
- Many elements, such as Gold, Silver, Iron, Phosphorus, Sulfur, Zinc, and Arsenic have been known since the pre-historic era.
- However, the first classification was made in the 18^{th} century.



18^{th} Century — Antoine Lavoisier work

- He attempted to classify known elements as metals and nonmetals.

Metals	Nonmetals



1829 — Dobereiner work

- In 1829, Döbereiner grouped the elements into triads (a group of three) with similar properties, noting that the atomic weight of the middle element was roughly the average of the other two.

Examples

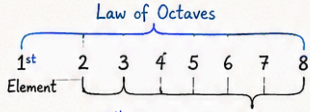
Such triads include lithium, sodium, and potassium (^{7}Li, ^{23}Na, ^{39}K).

$^{7}Li \quad ^{23}Na \quad ^{39}K$

$$\left(\frac{7 + 39}{2}\right) = 23$$

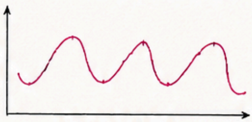
1864 — John Newlands work

- English chemist John Newlands, in 1864, first time observed periodicity in the 62 known elements were arranged in increasing order of their atomic masses. He classified the elements into groups so that every eight elements resembled the first element in properties.



1864 — Lothar Meyer work

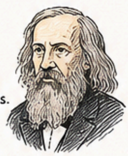
- In the same year, Lather Meyer developed his famous curves by plotting a graph between the atomic weight and atomic volumes of elements. These curves also showed periodicity.



Meyer's curves show periodic rise and fall in atomic volume.

1869 — Dmitri Mendeleev work

- In 1869, Russian chemist Dmitri Mendeleev, considered the father of the Periodic Table, arranged 63 elements by increasing atomic mass, aligning elements with similar properties into groups.
- The success of his table was hidden in leaving gaps for undiscovered elements and predicting their atomic mass and properties, which proved accurate when these elements were practically found.



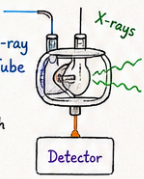
Mendeleev's Periodic Table (1869)

I	II	III	IV	V	VI	VII	VIII
H							
Li	Be	B	C	N	O	F	–
Na	Mg	Al	Si	P	S	Cl	–
K	Ca	–	Ti	Cr	Mn	Fe	Co Ni
–	Zn	–	–	As	Se	Br	–
–	–	–	In	Sn	Sb	Te	–
Cs	Ba	–	–	–	–	–	–

Gaps left for undiscovered elements

1913 — Moseley work

- In 1913, Moseley determined the exact atomic numbers of known elements using X-ray emission, resolving flaws and discrepancies in Mendeleev's table by arranging the elements by atomic numbers instead of atomic masses. This significant breakthrough led Moseley to modify the Periodic Law to state that the properties of elements are periodic functions of their atomic numbers.



Mendeleev (1869)	→	Moseley (1913)
Arranged by Atomic Mass (Flaws & Discrepancies)	→	Arranged by Atomic Number (Correct Order)

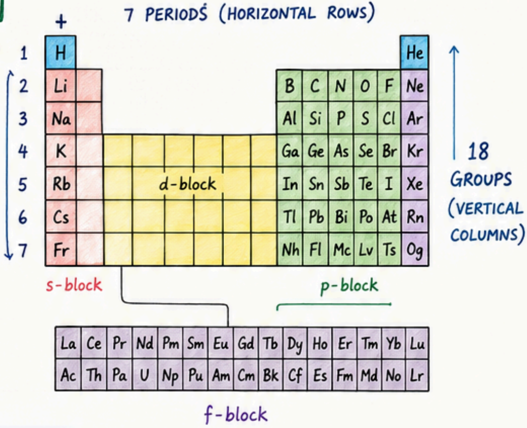
Modern Periodic Law
The properties of elements are periodic functions of their atomic numbers.

1.2 MODERN PERIODIC TABLE – MAIN FEATURES

The classification of elements in the modern periodic table helps in the easier understanding of their properties.

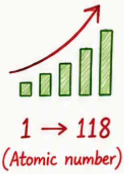
Features of modern periodic table

Following are some of the main features of the modern periodic table:



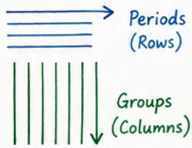
(ii) Ascending order of atomic number

Presently, 118 elements are grouped in the table in ascending order of their respective atomic numbers.



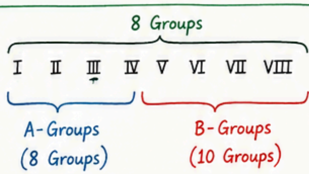
(iii) Periods and groups

There are seven horizontal rows called periods and eighteen vertical columns called groups.



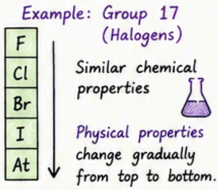
(iiii) Classification into sub-groups

In older versions of the table, there were 8 vertical groups, divided into two types of groups: Eight A-Groups and Ten B-Groups; sometimes.



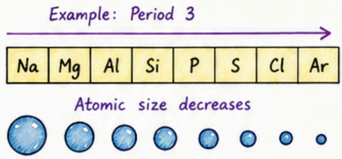
(iv) Properties based on group number

In the periodic table, elements within the same group exhibit similar chemical properties because they have the same number of valence electrons. However, they show a gradual change in physical properties from top to bottom in a group.



(v) Properties based on period number

Elements in a period show a gradual change in properties moving from left to right in periods. Atomic size gradually decreases as we move from left to right.



Remember!

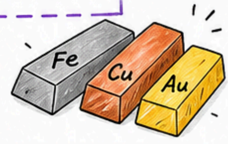
Down a group → change Top to bottom

Across a period → change Left to right

The modern periodic table is a powerful tool that helps us understand and predict the properties of elements in a systematic way.

1.3 METALS, NON-METALS AND METALLOIDS

Elements can be broadly classified as metals, nonmetals and metalloids.



1. METALS

- Metals are elements which tend to lose electrons to form positive ions.

Examples

Iron, copper, gold, silver, etc.

2. NON-METALS

- Non-metals are elements which tend to gain electrons to form negative ions.

Examples

Chlorine, sulphur, phosphorus, etc.

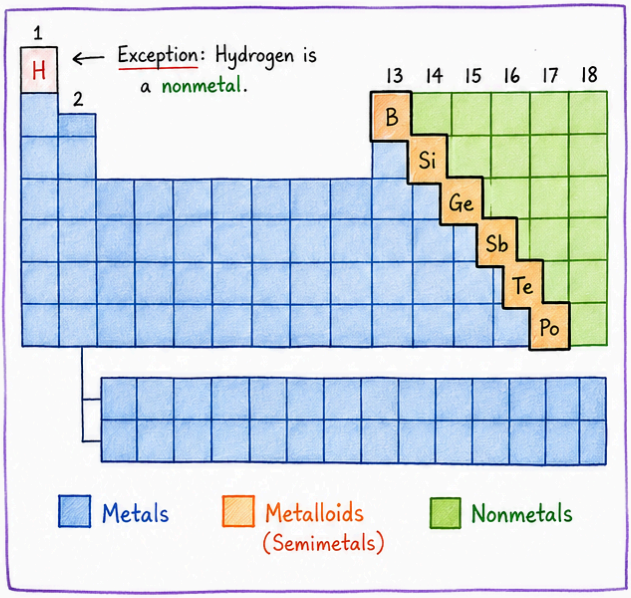
3. METALLOIDS

- "The metalloids exhibit some properties of metals and some of non-metals."

- The metalloids separate the metals and nonmetals on a periodic table.

STAIR-STEP LINE

1. Mostly periodic tables have a "stair-step line" on the table identifying the element groups.
2. The line begins at boron (B) and extends down to polonium (Po) including Si, Ge, As, Sb and Te.
3. Elements to the left of the line are considered metals.
4. Elements just to the right of the line exhibit properties of both metals and nonmetals and are termed as metalloids or semimetals.
5. Elements to the far right of the periodic table are nonmetals.
6. The exception is hydrogen, the first element on the periodic table.



SUMMARY

- Metals tend to lose electrons and form positive ions.
- Nonmetals tend to gain electrons and form negative ions.
- Metalloids show properties of both metals and nonmetals and lie along the stair-step line on the periodic table.

1:4 BLOCKS IN PERIODIC TABLE

Elements of periodic table are classified into four blocks on the basis of last electron present in the specific orbital.

(i) s-block elements

The valence electrons of elements in the first two groups are in the "s" subshells, placing these elements in the s-block.

Last electron in s subshell (ns^1 or ns^2)

(ii) p-block elements

The remaining elements in groups 13 to 18, including the inert gases in the last group, belong to the p-block.

Last electron in p subshell (np^1 to np^6)

(iii) d-block elements

Similarly, transition elements belong to the d-block.

Last electron in d subshell ($(n-1)d^1$ to $(n-1)d^{10}$)

(iv) f-block elements

The elements in the two series at the bottom of the table (known as Lanthanides and Actinides) are categorized as f-block elements.

Last electron in f subshell ($(n-2)f^1$ to $(n-2)f^{14}$)

Groups

1	2										13	14	15	16	17	18
H																He
Li	Be										B	C	N	O	F	Ne
Na	Mg										Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Cd	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Rg	Cn						

Lanthanides (4f series)	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinides (5f series)	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

s-block   d-block   p-block   f-block

NOTE:

Knowing which block an element belongs to reveals important information about its features, chemical reactivity, oxidation states, and other attributes such as electronegativity, ionization energy, electron filling, etc.

- ✓ Reactivity
- ✓ Oxidation states
- ✓ Electronegativity
- ✓ Ionization energy
- ✓ Electron filling
- ✓ etc.

1.5 Families in Periodic Table

☆ An element family is a set of elements sharing common properties.

☆ Elements may be categorized according to element families. There are five famous families of elements in the periodic table:

1. Alkali metals
2. Alkaline earth metals
3. Transition elements
4. Chalcogens
5. Halogens
6. Noble gases

(i) Alkali Metals

Group 1
3 Li
11 Na
19 K
37 Rb
55 Cs
81 Fr

Definition

"The elements of group 1 of the periodic table (Li, Na, K, Rb, Cs, Fr) except hydrogen are called as Alkali metals."

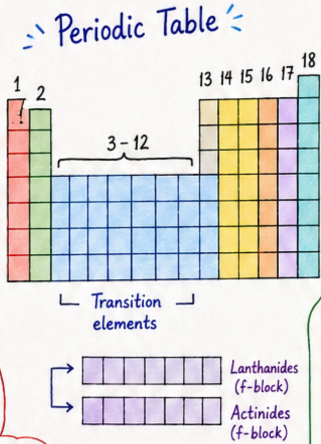
Properties

- They produce alkali in water.
- Alkali metals are characterized by having one valence electron.
- They are most reactive.

I-A group elements

Alkali metals include elements of Lithium ($_{31}Li$), Sodium ($_{111}Na$), Potassium ($_{19}K$), Rubidium ($_{37}Rb$), Cesium ($_{55}Cs$), and Francium ($_{81}Fr$).

Although hydrogen is not classified as an alkali metal due to its lack of typical group properties, it does share some properties with them.



(ii) Alkaline Earth Metals

Group 2
4 Be
12 Mg
20 Ca
38 Sr
56 Ba
88 Ra

Definition

"Group 2 elements (Be, Mg, Ca, Sr, Ba, Ra) are called Alkaline earth metals."

Group 2 elements

These include Beryllium ($_{14}Be$), Magnesium ($_{112}Mg$), Calcium ($_{200}Ca$), Strontium ($_{38}Sr$), Barium ($_{56}Ba$), and Radium ($_{88}Ra$).

General properties

- These metals are primarily found in the earth and form alkalis; hence they are called alkaline earth metals.
- These elements have two electrons in their valence shell.
- They are less reactive than alkali metals.

(iii) Transition Elements

The transition metals make up the largest family of elements in the middle of periodic table.

Example

They include four series of d-block elements, as well as lanthanides and actinides (f-blocks elements) found in the two rows below.

Examples include Cr, Fe, Co, Ni, Cu and Zn.

Properties

- They exhibit variable oxidation states.
- They mostly form coloured compounds.

(iv) Chalcogens

Group 16
8 O
16 S
34 Se
52 Te
84 Po
116 Lv

Groups 16 consists (O, S, Se, Te, Po, Lv) are called chalcogens.

Why called so?

The group 16 elements are called Chalcogens because most ores of copper (Greek chalkos) are oxides or sulfides.

Nature of elements

The chalcogens are composed of nonmetals O, S, metalloids Se, Te, Po and metals livermonium (Lv).

(v) Halogens

Group 17
9 F
17 Cl
35 Br
53 I
85 At
117 Ts

Elements in group 17 (F, Cl, Br, I, At, Ts), known as halogens.

Why called so?

The term "halogen" means "salt-former" because these elements easily react with alkali metals and alkaline earth metals to form stable halide salts.

Examples

Fluorine, Chlorine, Bromine, Iodine, Astatine and Tennessine.

General properties

- These are highly reactive nonmetals with high electron affinities.
- Halogens can easily accept one electron to complete their outermost shell.

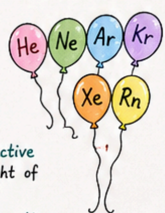
(vi) Noble Gases

Group 18
2 He
10 Ne
18 Ar
36 Kr
54 Xe
54 Xe
86 Rn
118 Og

"Elements in group 18 (He, Ne, Ar, Kr, Xe, Rn, Og) are called noble gases"

Examples

Helium, Neon, Argon, Krypton, Xenon and Radon.

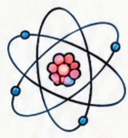


General properties

- The noble gases are a group of unreactive elements present at the extreme right of the periodic table in Group 18.
- Due to their stable electron configuration (complete outermost shell), they are almost entirely unreactive under normal conditions and rarely form compounds with other elements.

☆ Remember: Families help us predict properties of elements and understand chemical behaviour. ☆

1.6 PERIODIC ARRANGEMENT AND ELECTRONIC CONFIGURATION

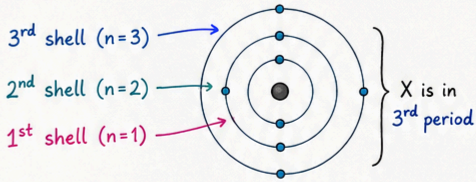


Understanding the periodic arrangement of elements in the periodic table offers valuable insight into their physical properties, such as their physical state and atomic radii, as well as their electronic structure and chemical reactivity.

Key Idea
Period → shells (n)
Group → valence electrons

① Relationship between period number and principal quantum number

The period number indicates the shell n, representing the number of electron shells surrounding the nucleus.

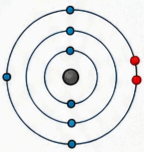


Example

- An element X in the 3^{rd} period has three electron shells, with its valence electrons located in the 3^{rd} shell.
- The specific subshell where the valence electrons are found, depends on the element's block subshell.
- If an element X in the 3^{rd} period is in the s-block, its valence electrons are in the 3s subshell.

② Relationship between group number and valence electrons

Additionally, the group number indicates the number of valence electrons.



3^{rd} period → 3 shells
Group 2 (s-block)
→ 2 valence electrons
Position : Period 3, Group 2
Element : Magnesium (Mg)

Example

An element X in the 3^{rd} period and group 2 has two valence electrons in its outermost shell. Thus, the element X in the 3^{rd} period and group 2 (s-block) has two valence electrons in the 3s subshell, which means that X would be magnesium (Mg).

③ Relationship between period and group number with electronic configuration

- Here is another example to relate period number and group number with electronic configuration and position of element in period table.

Example:

Element	Period	Group	Block	Valence e^-	Shells
X	3	13	p-block	3	3

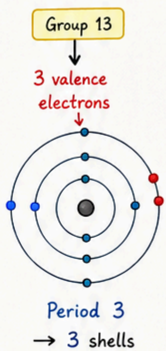
Electronic Configuration of X

$\underbrace{1s^2\ 2s^2\ 2p^6}_{\text{Inner shells}}\ \boxed{3s^2\ 3p^1}$

Valence electrons (in 3^{rd} shell) = 3

- In above example, the element X belongs to group 13 of periodic table so it has 3 valence electrons; and it is found in period 3 so it has three shells around its nucleus. It means that the 3-valence electron are in the 3^{rd} shell.
The configuration will be:

$\boxed{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1}$



Understanding the periodic arrangement of elements provides an explanation of an element's electronic configuration, which is essential for understanding its chemical properties and behavior.

1.7 PERIODICITY OF PROPERTIES

★ Modern Periodic Law

- **Statement**

> "The physical and chemical properties of elements are periodic functions of their atomic numbers."

▶ Studying the variation in properties with atomic numbers is important. We will study the following properties of elements in the periodic table.

We will study:

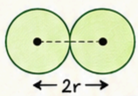
- Atomic Radius
- Ionization Enthalpy
- Electron Affinity
- Electronegativity
- Valency
- Metallic Character

1.7.1 ATOMIC RADIUS

▶ The atomic radius is a measure of the size of an atom.

Definition

"It is half of the distance between two identical atoms bonded together."



Units

The atomic radius is typically measured in picometers (pm) or Angstroms (Å).

$1\ pm = 10^{-12}\ m$ $\quad$ $1\ Å = 10^{-10}\ m$

Variation of atomic radius

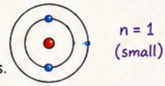
The atomic radius can vary depending on the type of bond (covalent, metallic & van der Waals forces) and the state of the atom. For example, the radius can be different in a covalent bond compared to an ionic bond.

Factors affecting the atomic radius

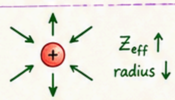
The factors affecting the atomic radius are:

(i) Number of shells
(ii) Effective nuclear charge and
(iii) Shielding effect of inner electrons.

(i) **Number of shells (n):** More shells means larger size because the outer electrons are farther from the nucleus.



(ii) **Effective nuclear charge (Z_{eff}):** Higher Z_{eff} pulls electrons closer, so the atomic radius decreases.



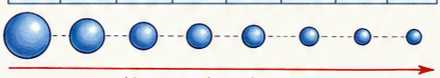
(iii) **Shielding effect:** More inner electrons shield the outer electrons from the nucleus, so radius increases.

PERIODIC TRENDS IN ATOMIC RADIUS

(a) Variation across the period

Generally, atomic radius decreases across a period (from left to right) in the periodic table due to increasing nuclear charge, which pulls the electron cloud closer.

3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne



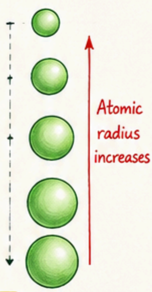
Atomic radius decreases

★ **Reason:** ↑ Nuclear charge → pulls electrons closer → smaller radius.

(b) Variation down the group

Atomic radius increases down a group (from top to bottom) because additional electron shells are added, so more shielding makes the atom larger despite the increase in nuclear charge (which is outweighed).

3 Li
11 Na
19 K
37 Rb
55 Cs



★ **Reason:** ↑ Number of shells & ↑ shielding effect → larger radius.

Key Takeaway: Atomic radius is not a fixed value; it changes systematically in the periodic table due to changes in number of shells, effective nuclear charge and shielding effect.

1.7.2 IONIC RADIUS

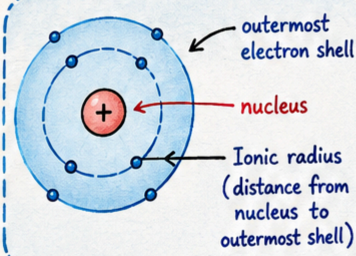
The ionic radius is a measure of the size of an ion in a crystal lattice.

DEFINITION

"It is defined as the distance from the nucleus of an ion to the outermost electron shell."

UNITS

It is measured in picometers (pm) or angstroms (Å).

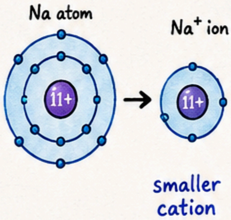


A. Smaller cationic radius than parent atom

When an atom loses one or more electrons to become a positive ion, it generally becomes smaller than the neutral atom.

Reason

This is because the loss of electrons reduces electronic repulsion and allows the remaining electrons to be pulled closer to the nucleus.

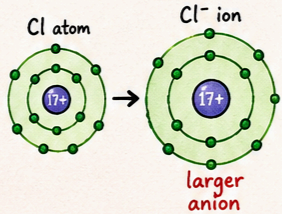


B. Larger anionic radius than parent atom

Contrarily when an atom gains one or more electrons to become an anion, it generally becomes larger than the neutral atom.

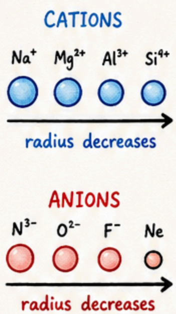
Reason

This is because the addition of electrons increases electronic repulsion, as a result the nuclear pull on electrons decreases and the electron cloud expands.



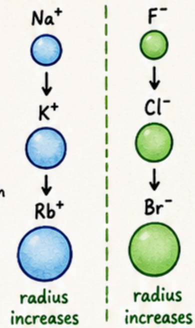
Variation across the period

- As you move across a period from left to right, the ionic radius of cations decreases due to the increasing nuclear charge which pulls the electrons closer.
- For anions, the ionic radius also decreases across a period because the increasing nuclear charge also pulls the electrons closer to the nucleus, even though anions are typically larger than the neutral atoms from which they are formed.



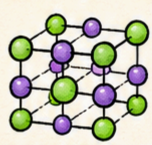
Variation down the group

- On the other hand, both cations and anions increase in size as we move down a group.
- This is because the principal quantum number (n) increases, leading to an increase in the number of electron shells.
- Consequently, the distance between the nucleus and the outermost electrons becomes larger, outweighing the effect of increased nuclear charge.
- The additional electron shells make the ions larger.



KEY TAKEAWAY

- Cations are smaller and anions are larger than their parent atoms.
- Ionic radius ↓ across a period (left → right)
- Ionic radius ↑ down a group (top → bottom)



1.7.3 Ionization Energy

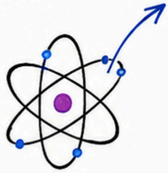
Applies to gaseous atoms and ions!

① First ionization energy

"The energy needed to remove one electron from each atom in one mole of atoms of the element in the gaseous state to form one mole of gaseous 1+ ions is known as 1^{st} ionization energy (ΔH_{i_1})."

$$Na_{(g)} \rightarrow Na^{+}_{(g)} + e^{-} \qquad \Delta H_{i_1} = 494 \text{ kJ mol}^{-1}$$

$$Ca_{(g)} \rightarrow Ca^{+}_{(g)} + e^{-} \qquad \Delta H_{i_1} = 590 \text{ kJ mol}^{-1}$$



e^{-}

② Second ionization energy

The removal of a second electron from each ion in a mole of gaseous 1+ ions is referred to as the second ionization energy (ΔH_{i_2}).

Again, with calcium as an example:

$$Ca^{+}_{(g)} \rightarrow Ca^{2+}_{(g)} + e^{-} \qquad \Delta H_{i_2} = 1150 \text{ kJ mol}^{-1}$$

We remove the 2^{nd} electron from a 1+ ion.

③ Third ionization energy

The removal of a third electron from each ion in a mole of gaseous 2+ ions corresponds to the third ionization energy (ΔH_{i_3}).

Again, with calcium as an example:

$$Ca^{2+}_{(g)} \rightarrow Ca^{3+}_{(g)} + e^{-} \qquad \Delta H_{i_3} = 4940 \text{ kJ mol}^{-1}$$

We remove the 3^{rd} electron from a 2+ ion.

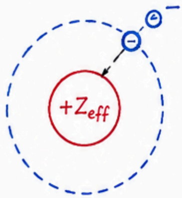
NOTE: An element can have several ionization energies; the exact number corresponds to its atomic number.

Factors affecting the ionization energy

The magnitude of the ionization energy of an element depends upon the following factors :

(i) Nuclear charge

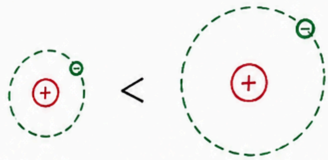
Greater the effective nuclear charge, greater is the electrostatic force of attraction, more difficult is the removal of an electron from the atom. For this reason, ionization energy increases with an increase in the effective nuclear charge.



Greater Z_{eff}
⇒ stronger attraction
⇒ more difficult to remove electron
⇒ higher ionization energy

(ii) Size of the atom or ion

In bigger atoms force of attraction between the nucleus and the outermost electrons is weaker. Therefore, the ionization energy decreases as the size of the atom increases and vice-versa.



Smaller atom (stronger attraction) ⇒ higher I.E.

Larger atom (weaker attraction) ⇒ lower I.E.

(iii) Electronic arrangement

It is observed half-filled and completely-filled orbitals are found to be more stable. Therefore, the ionization energy is higher when an electron is to be removed from a fully filled or half-filled-subshell.

Noble gas (ns^2np^6)
⇒ completely-filled orbital
⇒ highest ionization energy in the period

(a) Noble gases have highest ionization energies in their respective periods. It is due to highly stable fully-filled orbital (ns^2np^6).

	2s	2p			
N :	↑↓	↑	↑	↑	(Half-filled subshell)
O :	↑↓	↑	↑		(Not half-filled less stable)

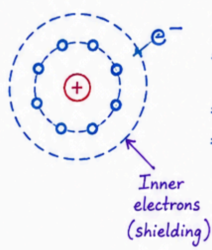
(b) Oxygen has lower ionization energy than nitrogen. The electronic configuration of oxygen and nitrogen are

$_7N = 1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ (more stable) $= 1403\ kJ\ mol^{-1}$

$_8O = 1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^0$ (less stable) $= 1365\ kJ\ mol^{-1}$

Shielding effect

Greater the shielding, easier it is to remove the valence electrons from an atom. Larger the number of inner electrons, greater is the screening effect, therefore, lower is the ionization energy.



More shielding
⇒ weaker attraction on valence electron
⇒ easier to remove
⇒ lower ionization energy

(v) SPIN-PAIR REPULSION

When electrons are spin-paired in the same orbital, the repulsion between them can lead to a slightly lower ionization energy compared to removing an unpaired electron. This is because the paired electrons experience increased repulsion, making it slightly easier to remove one of the paired electrons.

- Manganese (Mn) has two spin-paired electrons in its 4s orbital. The ionization energy to remove one of these paired electrons is relatively lower due to the increased repulsion between the paired electrons.

$\boxed{\text{Mn (25) : [Ar] } 3d^5\, 4s^2 \quad \text{(more stable)}}$

4s orbital : $\boxed{\uparrow\downarrow}$ Two paired electrons (↑↓) in 4s orbital
→ more repulsion
→ lower ionization energy

- In contrast, Chromium (Cr) has one unpaired electron in its 4s orbital. Removing one of these unpaired electrons requires more energy due to the absence of spin-pairing repulsion.

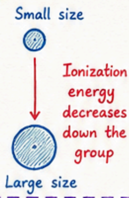
$\boxed{\text{Cr (24) : [Ar] } 3d^5\, 4s^1 \quad \text{(less stable)}}$

4s orbital : $\boxed{\uparrow}$ One unpaired electron (↑)
→ no spin-pair repulsion
→ higher ionization energy

PERIODIC TRENDS IN IONIZATION ENERGY

(a) Variation down the group

Going down in a group, ionization energy decreases from top to bottom due to increase in the atomic size and shielding effect.

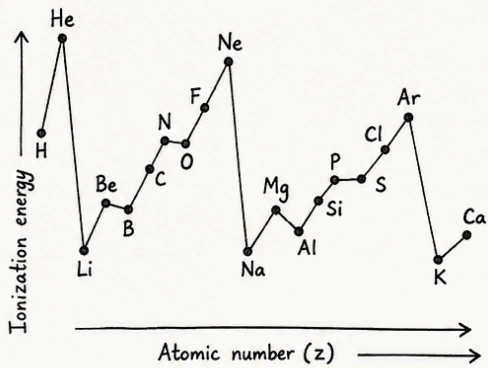


(b) Variation across the period

As you move from left to right across a period, the number of shells remains unchanged while the effective nuclear charge increases. Therefore, ionization energy increases from left to right.

- Although the number of electrons also increases across a period, the shielding effect within the same shell is same so not considered.
- The trend of ionization energies of period (1–3) is shown in Fig 1.5.

The figure also reveals that Noble gases have the highest values of ionization energy because due to complete outermost shell in them, the removal of electron is extremely difficult, whereas alkali metals have lowest values of ionization energy.



Alkali metals (Li, Na, K)	Noble gases (He, Ne, Ar)	General trend
↓ Lowest ionization energy (one valence electron, easy to remove)	↓ Highest ionization energy (complete outermost shell, very difficult to remove an electron)	Increases from left to right across a period due to increase in effective nuclear charge.

Fig 1.5 Variation of ionization energies across periods

1.7.4 Electron Affinity (ΔH_{ea}°)

First electron affinity

- "The first electron affinity, ($\Delta H_{ea_1}^\circ$), is the enthalpy change involved when 1 mole of electrons is added to 1 mole of gaseous atoms to form 1 mole of gaseous uni-negative ions under standard conditions."

Example

Electron affinity of chlorine atom,

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \qquad \Delta H_{ea_1}^\circ = -348.8\ \text{kJ mol}^{-1}$$

This is amount of energy released when 6.02×10^{23} atoms of chlorine in the gaseous state are converted into $Cl^-_{(g)}$ ions. Since, energy is released, so first electron affinity carries negative sign.

Second electron affinity

The second electron affinity, $\Delta H_{ea_2}^\circ$ is the amount of energy required to add electrons to 1 mole of uni-negative gaseous ions to form 1 mole of gaseous 2− ions under standard conditions.

Example

When first electron is added to a neutral oxygen atom, 141 kJ mol^{-1} energy is released.

$$O_{(g)} + e^- \longrightarrow O^-_{(g)} \qquad \Delta H_{ea_1}^\circ = -141\ \text{kJ mol}^{-1}$$

But 798 kJ mol^{-1} of energy is absorbed on adding second electron to a uni-negative (O^-) ion.

$$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} \qquad \Delta H_{ea_2}^\circ = +798\ \text{kJ mol}^{-1}$$

Factors affecting electron affinity

Important factors affecting the magnitude of electron affinity values of elements are as follows:

(i) Size of atom

For small sized atoms the attraction of the nucleus for the incoming electron is stronger. Thus, smaller is the size of the atom, greater is its electron affinity.

(ii) Nuclear charge

Greater the magnitude of nuclear charge of an element stronger is the attraction of its nucleus for the incoming electron. Thus, with the increase in the magnitude of nuclear charge, electron affinity also increases.

(iii) Electronic configuration of atom

The electron affinity is low when the electron is added to a half-filled sub-shell than that for partially filled one.

This can be explained by considering the following examples.

Electron affinity values of 'N' and 'P' group-15 (V-A), atoms are very low.

Reason

This is because of the presence of half-filled 'np' orbitals in their valence shell ($N = 2s^2\ 2p^3$, $P = 3s^2\ 3p^3$). These half-filled p-subshells, being very stable, have very little tendency to accept any extra electron to be added to them.

N (Z = 7)
$2s^2\ 2p^3$

2s		2p	
↑↓	↑	↑	↑

P (Z = 15)
$3s^2\ 3p^3$

3s		3p	
↑↓	↑	↑	↑

(Half-filled p-subshells)

Periodic trends in electron affinity

(a) Variation down the group

- As the atomic size increases down the group, the larger electron cloud causes the incoming electron to experience less attraction from the nucleus.
- Consequently, electron affinity generally decreases down the group.
- This trend is observed in the halogens

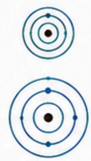
$$At < I < Br < F < Cl$$

Group 17 (Halogens)	Electron affinity ($kJ\ mol^{-1}$)
Cl	− 349
F	− 328
Br	− 325
I	− 295
At	− 270

decreases down the group

Down the group

Atomic size increases



Attraction on incoming electron decreases

(b) Variation across the period

- Generally, electron affinities become more negative as we move from left to across a period.

Across a period (left ⟶ right)

Li	Be	B	C	N	O	F
−60	≥0	−27	−122	≥0	−141	−328

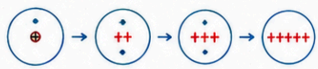
Electron affinity ($kJ\ mol^{-1}$)

More negative electron affinity (greater tendency to accept electron)

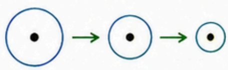
This is firstly due to increase in the nuclear charge, which attracts additional electrons more strongly and secondly due to decreasing atomic radius.

Factors responsible

① Nuclear charge increases
(More protons ⟶ stronger attraction for incoming electron)

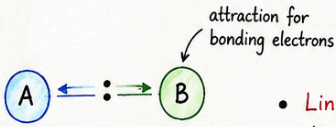


② Atomic radius decreases
(Electrons are closer to the nucleus)



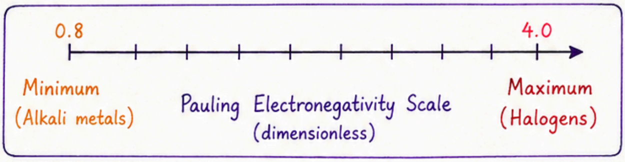
1.7.5 Electronegativity

Definition



"Electronegativity is a measure of an atom's attraction for bonding electrons in a molecule relative to other atoms."

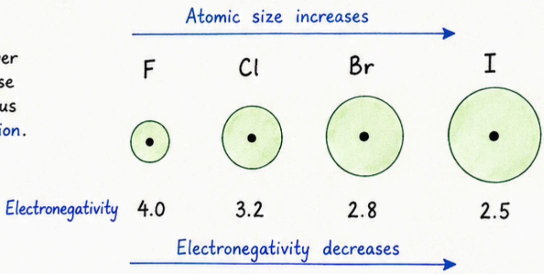
- Linus Pauling, an American chemist, developed a scale of dimensionless electronegativity values.
- on this scale, alkali metals have minimum electronegativity (0.8), and halogens have maximum values (F = 4.0).



Factors Affecting Electronegativity

(i) Atomic size

A larger atomic size will result in a lower value of electronegativity. This is because electrons being far away from the nucleus will experience a weaker force of attraction.



Example (Group 17 – Halogens)

The electronegativities of halogens in group 17 are in the order

F > Cl > Br > I

It is due to the increase in the atomic size from F to I in the group.

(ii) Effective nuclear Charge

A higher value of the effective nuclear charge will result in a greater value of electronegativity, because an increase in nuclear charge causes greater attraction to the bonded electrons.

This is why the electronegativity in a period increases from left to right.

Example (Period 2)

	Li	Be	B	C	N	O	F
Electronegativity (Pauling scale)	1.0	1.5	2.0	2.5	3.0	3.5	4.0

Electronegativity increases →

Electronegativity of Li = 1.0
Electronegativity of F = 4.0

Periodic Trends in Electronegativity

Variation across the period

- When we move from left to right along the period, the electronegativity increases. This is due to increasing nuclear charge and decreasing size.
- Normally, metals being on the left side of the periodic table, possess lower electronegativity values than those of non-metals.

Hence, metals are electropositive and non-metals are electronegative, relatively.

Variation down the group

- In the groups, it decreases from top to bottom. This is due to the increase in size by the addition of shells and increasing shielding effect.
- For example, in the halogen group, the electronegativity value decreases from fluorine (4.0) to iodine (2.5) as shown in a part of the table in Figure 1.6.

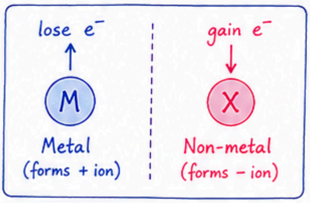
Figure 1.6 (Part)

Group 17 (Halogens)			Electronegativity
9	F	Fluorine	4.0
17	Cl	Chlorine	3.2
35	Br	Bromine	2.8
53	I	Iodine	2.5

Electronegativity decreases from top to bottom.

1.7.6 Variation in Metallic Character

The metallic character of elements is typically their tendency to lose electrons.



- We find that elements on the left side of the periodic table have a greater tendency to lose their outermost electrons to achieve a complete octet are metals and produce positive ions.
- In contrast, elements on the right side of the table tend to gain electrons to complete their outermost shell are non-metals and produce negative ions.

Conclusion → Hence one can conclude that the metallic character of an element largely depends on its valence shell electronic configuration.

Variation across the period

- Across the period (from left to right), we find that the outermost shell remains the same, but the number of protons (positive charge) in the nucleus increases.
- Consequently, the metallic character of the elements decreases.
- In other words, the increase in nuclear charge pulls the electron cloud closer to the nucleus, making it more difficult for the atom to lose electrons and thereby decreasing its metallic character.

Thus, metallic character decreases across a period from left to right.

Variation down the group

- Contrarily metallic character increases as one moves down in a group of the periodic table. This is due to the fact that the electrons become easier to lose as the atomic size (radius) increases.
- As the atomic size increases, the electrostatic force of attraction decreases between the nucleus and the electrons, causing the electrons to be held more loosely.
- The increase in metallic character (ease of losing electron) makes the element more reactive.

☆ **Conclusion** → Cesium is far more reactive and electropositive than sodium or lithium.

$\text{Li} < \text{Na} < \text{Cs}$

Metallic character increases down the group.

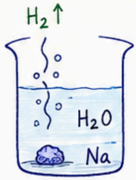
(Easier to lose electrons)

1.8 REACTIONS OF SODIUM AND MAGNESIUM

1.8.1 With water

- Sodium is more reactive than magnesium towards water. Na reacts vigorously with water to form sodium hydroxide and hydrogen.

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

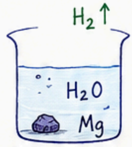


Vigorous reaction with water!

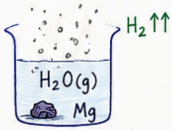
- Mg reacts more slowly in forming magnesium hydroxide and hydrogen. However, magnesium reacts with steam more vigorously to make magnesium oxide and hydrogen gas.

$$\boxed{Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)} \leftarrow \text{With water}$$

$$\boxed{Mg(s) + 2H_2O(g) \longrightarrow MgO(s) + 2H_2(g)} \leftarrow \text{With steam}$$



Slow reaction



Vigorous with steam!

1.8.2 With Oxygen

- Sodium burns in oxygen with an golden yellow flame to produce a white solid mixture of sodium oxide and sodium peroxide. Sodium is kept under kerosene oil to prevent its reaction with air. It reacts vigorously with oxygen in open air to form peroxide.

Peroxide (in excess O_2)

$$\boxed{2Na(s) + O_2(g) \longrightarrow 2Na_2O_2(s)}$$

Under special conditions like limited O_2 or high temperature, sodium oxide is formed.

Oxide (limited O_2 / high temp.)

$$\boxed{4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)}$$

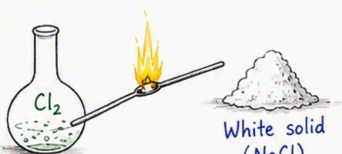
- Magnesium burns in oxygen with an intense white flame to give white solid magnesium oxide.

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

1.8.3 With Chlorine

- Chlorine reacts with both metals to give soluble salts. It reacts exothermically with sodium, golden yellow flame is seen and white solid, sodium chloride is formed.

$$\boxed{2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)}$$



White solid (NaCl)

- Magnesium also reacts with chlorine to give white solid, magnesium chloride.

$$\boxed{Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)}$$

White solid ($MgCl_2$)

1.9 TRENDS IN BONDING IN OXIDES AND CHLORIDES OF PERIOD 3

Oxides of period 3

- Oxides of group 1, 2 & 3 (e.g., Na_2O) have more ionic character. These oxides exist as giant ionic lattices with strong electrostatic forces between oppositely charged ions.
- Oxides of group 4, 5, 6 & 7 (e.g., SO_2) are more covalent. These oxides exist as covalent molecules with weak intermolecular forces between molecules.
- This transition is a result of the increasing electronegativity and decreasing ionic character.

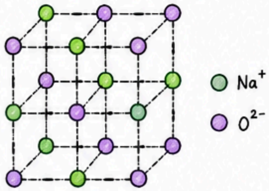
Period 3 elements →

Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7
Na	Mg	Al	Si	P	S	Cl

More ionic (character) ⟷ More covalent (character)

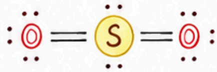
e.g., Na_2O

Giant ionic lattice



e.g., SO_2

Covalent molecules



Weak intermolecular forces between molecules

Chlorides of period 3

- Similar to oxides, chlorides of group 1, 2 and 3 (e.g., NaCl) are predominately ionic.
- Chlorides of elements from group 4, 5, 6 and 7 (e.g., PCl_5) are covalent.
- The covalent character in chlorides increases due to decrease in difference of electronegativity between the halogen and the other atom the higher electronegativity of the central atom.

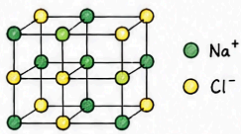
← Period 3 elements →

Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7
Na	Mg	Al	Si	P	S	Cl

More ionic (character) ⟷ More covalent (character)

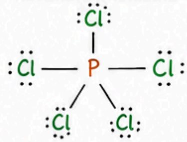
e.g., NaCl

Ionic lattice



e.g., PCl_5

Covalent molecule



1.9.1 CLASSIFICATION OF OXIDES

Oxides

- Oxides are binary compounds formed by the reaction of oxygen with other elements.
- Oxygen is highly reactive in nature hence it reacts with metals and non-metals to form oxides.
- The classification of oxides is done into neutral, amphoteric and basic or acidic based on their acid-base characteristics. Along the period three nature of oxides will be explored.

1. BASIC OXIDES

"A basic oxide is an oxide that when combined with water gives off an alkali."

Example

Na_2O, CaO, BaO etc.

NOTE:

Solubility group 2 hydroxides increases down the group so alkalinity also increases down the group.

Properties

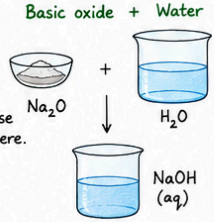
(ii) Metals react with oxygen to give basic oxides.

(iii) These oxides are usually ionic in nature.

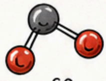
(iiii) Group 1 and 2 form basic oxides when react with oxygen. These compounds readily react with water but few exceptions are there.

$$Na_2O_{(s)} + H_2O(\ell) \rightarrow 2NaOH(aq)$$
$$CaO_{(s)} + H_2O(\ell) \rightarrow Ca(OH)_2(aq)$$

Basic oxide + Water



2. Acidic oxides



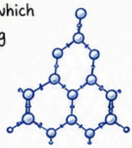
SO_2

"An acidic oxide is an oxide that when combined with water gives off an acid.

Properties

(i) Non-metals react with oxygen to form acidic oxides which are held together by covalent bonds. These compounds can also be called acid anhydrides.

(ii) Acidic oxides usually have a low melting and boiling point except for oxides like SiO_2 which is insoluble in water and have high melting points as it forms giant molecules.



SiO_2
(giant molecule)

Examples

- Examples of acidic oxides in period 3 are: P_2O_3, P_2O_5, SO_3, SO_2.
- Silicon dioxide is acidic oxide as it can react with bases.

Reactions with Water

- $P_2O_3(s) + H_2O(\ell) \rightarrow H_3PO_3(aq)$
- $P_2O_5(s) + H_2O(\ell) \rightarrow H_3PO_4(aq)$
- $SO_2(g) + H_2O(\ell) \rightarrow H_2SO_3(aq)$
- $SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(aq)$

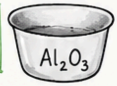
Reactions with Bases

Reactions of theses oxides with bases are given below:

- $SiO_2(s) + 2NaOH(aq) \rightarrow Na_2SiO_3(aq) + H_2O(\ell)$
- $P_2O_3(s) + 6NaOH(aq) \rightarrow 2Na_3PO_3(aq) + 3H_2O(\ell)$
- $SO_2(g) + 2NaOH(aq) \rightarrow Na_2SO_3(aq) + H_2O(\ell)$

3. Amphoteric oxides

"Amphoteric oxides are oxides that can react with both acids and bases."



Properties

(i) They have the ability to behave as either an acid or a base, depending on the conditions.

(ii) In chemical reactions, they can neutralize both acidic and basic substances.

(iii) This dual reactivity is a characteristic feature of amphoteric oxides, distinguishing them from other oxides that are typically either basic or acidic.

Example

Aluminum oxide (Al_2O_3) is insoluble in water but reacts with hydrochloric acid to form aluminium chloride and water, and with sodium hydroxide to form sodium aluminate and water.

$$\underset{\text{acid}}{Al_2O_3(s) + 6HCl(aq)} \rightarrow \underset{\text{salt}}{2AlCl_3(s)} + 3H_2O(\ell)$$

$$\underset{\text{base}}{Al_2O_3(s)} + \underset{\text{base}}{2NaOH(aq)} \rightarrow \underset{\text{sodium aluminate}}{Na_2Al_2O_4(s)} + H_2O(\ell)$$

1.9.2 Classification of Chlorides

"Chlorine is a highly reactive nonmetal that forms stable compounds known as chlorides through chemical reactions with various elements."

One common example is table salt, an ionic compound consisting of sodium and chloride. These chlorides show characteristic behavior when we add them into water, resulting in solutions that can be acidic or neutral.

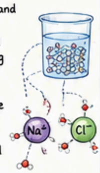
1. Neutral chlorides

"Neutral chlorides are salts that, when dissolved in water, produce a neutral solution with a pH close to 7."

Solubility in Water

At the start of period 3, chloride sodium and magnesium do not react with water.

- The polar water molecules are attracted to the oppositely charged ions, dissolving these chlorides by breaking down their giant ionic lattices.
- The solutions formed contain the positive metal ions and negative chloride ions surrounded by water molecules. These ions are now known as hydrated ions and this process is known as hydration.



Example

When sodium chloride (NaCl) is dissolved in water, it undergoes dissociation into its constituent ions.

$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ (pH = 7)

$MgCl_2(s) \rightarrow Mg^{+2}(aq) + 2Cl^-(aq)$ (pH = 6.5)

NOTE: Group 1 chlorides are all neutral chlorides and group 2 chlorides are also neutral with few exceptions.

2. Acidic chlorides

"The chlorides react with water to make acidic solution with pH less than 7 are called acidic chlorides."

- If we move in period 3, from aluminum to sulphur all chlorides react with water to make acidic solution with pH less than 7 this process is called hydrolysis.

Example

Aluminium chloride exist as dimer Al_2Cl_6 which is covalently bonded. Once we add water, dimer breaks and aluminium and chloride ions in solution. Al^{3+} ion is hydrated and causes a water molecule to lose an H^+ ion, this process is hydrolysis. This turns the solution acidic.

The following reactions shall occur:

$$AlCl_3(s) + 3H_2O(\ell) \rightarrow \underset{\text{aluminium hydroxide (precipitate)}}{Al(OH)_3(s)} + \underset{\text{acidic solution}}{3HCl(aq)}$$

- Other examples of acidic chlorides are given below.

$$SiCl_4(\ell) + 2H_2O(\ell) \rightarrow \underset{\text{silicon dioxide (precipitate)}}{SiO_2(s)} + \underset{\text{acidic solution}}{4HCl(aq)}$$

SiO_2

$$PCl_3(\ell) + 3H_2O(\ell) \rightarrow \underset{\text{phosphorous acid (precipitate)}}{H_3PO_3(s)} + \underset{\text{acidic solution}}{3HCl(aq)}$$

1.10 VARIATION IN OXIDATION NUMBER IN OXIDES AND CHLORIDES

- The oxidation number of an atom is the formal charge on that atom in a molecule or ion.
- The oxidation number is also referred to as the oxidation state.
- In ionic compounds the oxidation number of an atom is defined as the charge which appears on the ions.

Let's examine the oxidation numbers in oxides and chlorides of the third period.

Oxidation number of 3rd period elements

The oxidation number of an element of 3rd period in its oxide or chloride corresponds to the number of electrons used for bonding and is always positive because oxygen and chlorine are more electronegative than any of these elements.

Oxidation state and group number

The maximum oxidation number matches the group number, reflecting the total number of valence electrons.

Consider the following table (Table 1.2) for oxidation states of various elements of the periodic table.

In the oxides, the maximum oxidation number increases from +1 in Na to +6 in S.

Element (3rd period)	Na (Group 1)	Mg (Group 2)	Al (Group 13)	Si (Group 14)	P (Group 15)	S (Group 16)	Cl (Group 17)
Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	—
Maximum oxidation number in oxides	+1	+2	+3	+4	+5	+6	—
Chloride	$NaCl$	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_5	SCl_2	Cl_2
Maximum oxidation number in chlorides	+1	+2	+3	+4	+5	+2	+1

Oxidation state in chlorides

In chlorides, the maximum oxidation number increases from +1 in Na to +5 in P.

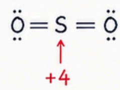
$$\begin{array}{ccccccccc} Na & & Mg & & Al & & Si & & P \\ +1 & \rightarrow & +2 & \rightarrow & +3 & \rightarrow & +4 & \rightarrow & +5 \end{array}$$

Oxidation state in phosphorus and sulfur

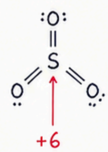
Phosphorus and sulfur exhibit several oxidation numbers because they can expand their octet by exciting electrons into empty 3d orbitals.

Example

- In SO_2, sulfur has an oxidation number of +4 because only four electrons are used for bonding.



- In SO_3, sulfur has an oxidation number of +6 because all six electrons are used for bonding.

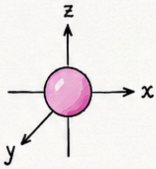


Quick Check 1.1

(a) Why are the elements in Groups 1 and 2 known as s-block elements?

Ans. Group 1 and 2 elements are called s-block elements because outermost electrons of their atoms are in s-orbital.

s-orbital (s)



s-orbital is spherically symmetric.

Max. electrons in s-orbital = 2

(b) Name the elements in the chalcogen family. Give their two characteristics.

Ans. Groups 16 consists (O, S, Se, Te, Po, Lv) are called chalcogens.

Group 16 = Chalcogens (6 valence e^-)

Names	Oxygen (O), Sulfur (S), Selenium (Se), Tellurium (Te), Polonium (Po), Livermonium (Lv).
Properties/ Characteristics	(i) They all have six valence electrons.

Quick Check 1.2

(a) X belongs to group 14 and period 2.

i. Write electronic configuration of the element X.

ii. Identify block of the element. Identify this element from periodic table.

Group 14 elements have 4 valence electrons ($ns^2\ np^2$)

Ans. i. Electronic configuration of X:

$\boxed{1s^2\ \ 2s^2\ \ 2p^2}$

[↑↓] 2s [↑↓] 2p [↑ | ↑ |]

Period 2

13	14	15	16	17	18
B Boron	C Carbon	N Nitrogen	O Oxygen	F Fluorine	Ne Neon

Group 14

ii. X = it is an element of p-block which is carbon and belongs to group 14 and period 2.

(b) Predict the electronic configuration of an element in period 4 and Group 17 without consulting the periodic table.

Ans. Group 17 elements are halogens having 7 valence electron (ns^2, np^5) and period 4 means ($n = 4$), so valence electron configuration should be $4s^2$, $4p^5$ and final configuration will be

Group 17 = Halogens (7 valence e^-) ($ns^2\ np^5$)

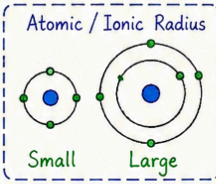
$$_{35}Br = \underbrace{1s^2\ 2s^2\ 2p^6}_{n=1}\ \underbrace{3s^2\ 3p^6\ 3d^{10}}_{n=3}\ \underbrace{4s^2\ 4p^5}_{n=4}$$

Valence shell ($n = 4$) has $4s^2\ 4p^5$ (7 valence e^-)

Quick Check 1.3

(a) Which factors affect atomic and ionic radii?

Ans. There are factors which affects the atomic and ionic radii. Let us discuss them one one.



(i) Effective nuclear changes

As the effective nuclear change increases the force of attraction of nucleus increases on the valance shell electrons and size of atom and ion decreases.

(ii) Shielding effects

With the increase in shielding effect the nuclear force of attraction on valance electrons decreases which cause on increase in atomic and ionic radii.

(b) Using your knowledge of Period 3 elements, predict and explain the relative sizes of

i. The atomic radii of lithium and fluorine
ii. A lithium atom and its ion, Li^+
iii. An oxygen atom and its ion, O^{2-}
iv. A nitride ion, N^{3-}, and a fluoride ion, F^-

Period 3 elements

Na Mg Al Si P S Cl Ar

Atomic size decreases ⟶

Ans.

i. The atomic radius of Li is greater than F because when we move from left to right the atomic size decreases. $Li > F$

ii. The size of Li atom is greater than that of Li^+ because when the electron removed from Li to convert it into Li^+ the 2nd shell is removed so the size is decreased. $Li > Li^+$

iii. O^{2-} has greater size than oxygen atom because by gaining electron, the repulsions between valence electron increases and size increases. $O^{2-} > O$

iv. N^{3-} has greater size than F^-
N^{3-} and F^- are isoelectronic but N^{3-} has more -ve charge on it and has maximum electronic repulsion which increases its size. $N^{3-} > F^-$

Quick Check 1.4

Ionization Energy (IE)

Energy required to remove the most loosely bound electron from a gaseous atom.

(a) Explain with reasoning following facts about ionization energy:

i. 1st ionization energy of Boron is lesser than Beryllium.
ii. 1st ionization energy of Aluminium is lower than Magnesium.

Ans.

i. 1st ionization energy of boron is lesser than beryllium because of the electronic configuration of Be. The Be has complete s-subshell due to which more energy is required to remove electron which B has partially filled p-subshell.

Be : $1s^2$ $2s^2$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ B : $1s^2$ $2s^2$ $2p^1$ $\boxed{\uparrow}\boxed{}\boxed{}\boxed{}$

ii. Al is present under B and Mg is present under Be. Mg also have completely filled s-subshell and Al has partially filled p-subshell so, that the 1st ionization energy of Mg is greater the Al.

Reminder!

Across a period IE generally increases. →

Down a group IE generally decreases. ↓

(b) What trend is observed in ionization energy as you go down group 3? Give reason.

Ans. Trend in ionization energy

Moving down the group the number of shells increases due to which atomic size increases, Z_{eff}. decreases, shielding effect increases which cause decreases in ionizations energy as less amount of energy is required to remove electron from the atom.

Group 3

B (2 shells)
Al (3 shells)
Ga (4 shells)
In (5 shells)
Tl (6 shells)

IE decreases down the group ↓

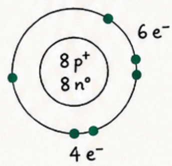
Quick Check 1.5

Explain with reasoning following facts about electron affinity :

(a) 1st electron affinity of Oxygen is −141 kJ/mol but 2nd electron affinity is +844.0 kJ/mol.

Ans. 2nd electron affinity of oxygen is lesser than the 1st because as the 1st electron enters in the valance shell it causes repulsions to increase between valance shell electrons now the second entering electron face more difficulty to enter in the valance shell hence electron affinity decreases.

O (Z = 8)



1st e⁻ enters → repulsions increase
↓
2nd e⁻ faces more difficulty
↓
EA decreases (+ve value)

(b) Which of nitrogen and phosphorus has the higher electron affinity? Justify with reason.

Ans. Nitrogen has greater electron affinity as comparand to phosphorus because P is present under N in the same group. We know that from top to down the electron affinity decreases generally.

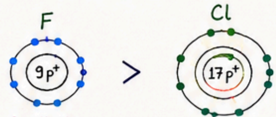
Group 15

N
↓
P
↓
As
↓
Sb
↓
Bi

Electron affinity decreases down the group

(c) F has lower electron affinity than Cl although its size is smaller. Explain why?

Ans. F has very small size and dense electronic cloud so in coming electron is repelled consequently F has less electron affinity than that of Cl.

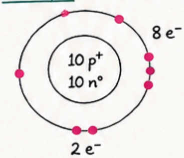


F	Cl
Small size	Larger size
Dense e⁻ cloud	Less dense e⁻ cloud
More repulsion	Less repulsion
Lower EA	Higher EA

(d) Why noble gases (group-18) have positive 1st electron affinities? Explain in terms of electronic configuration

Ans. Nobel gases have completely filled valance shell so adding an electron makes them less stable and energy is required to add an electron.

Example : Ne (Z = 10)



Complete octet (very stable)
↓
Adding e⁻ breaks stability
↓
Energy required
⇒ EA is +ve

Quick Check 1.6

(a) Illustrate how does the metallic character vary in group 14.

Ans. In group 14 elements, metallic character increases as you move down the group. This trend is due to increasing atomic size and decreasing ionization energy, making it easier for elements to remove electrons and form positive ions.

Group 14

C
↓
Si
↓
Ge
↓
Sn
↓
Pb

- Non-metal = C
- Metalloid = Si and Ge
- Metals = Sn and Pb

Trend down the group:

- Atomic size ↑
- Ionization energy ↓
- Electrons are lost more easily
- Metallic character ↑

Metallic character increases down the group

(b) Identify semi metals in groups 14, 15 and 16. Why they are semi metals?

Ans. Ge in 14 Group, As, Sb in 15-Group and Te in 16-Group are semi-metals and are called so because they have properties between metals and non-metals.

Group 14 Semi-metal	Group 15 Semi-metals	Group 16 Semi-metal
Ge (Germanium)	As (Arsenic) Sb (Antimony)	Te (Tellurium)

They have intermediate properties between metals and non-metals. ★

QUICK CHECK 1.7

(a) What is the nature of oxides and hydroxides of Na and Mg?

Ans. **Oxides of Na and Mg**

Na_2O and MgO are basic in nature, these oxides produce alkalies when dissolved in H_2O.

$$Na_2O + H_2O \rightarrow 2NaOH$$
$$MgO + H_2O \rightarrow Mg(OH)_2$$

Na_2O and MgO both are ionic in nature.

Hydroxides of Na and Mg

$NaOH$ and $Mg(OH)_2$ are basic in nature, these hydroxides react with acids as follows:

$$NaOH + HCl \rightarrow NaCl + H_2O$$
$$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + H_2O$$

(b) What could you predict about the reactivity of Ca, a group 2 element, when reacted with water and oxygen?

Ans. Calcium reacts with oxygen slowly in air forming protective layer of CaO at room temperature when initialed calcium burns vigorously forming red flame and CaO (Quick time).

$$Ca_{(s)} + O_{2(g)} \xrightarrow[\text{(slow)}]{} CaO_{(s)}$$

When Ca metals reacts with H_2O it produces $Ca(OH)_2$ and H_2. The reaction is slow.

Note: Ca is more reactive than Mg but less reactive than group 1 metals.

QUICK CHECK 1.8

(a) ZnO reacts with HCl to give $ZnCl_2$ and with NaOH to give Na_2ZnO. Give equations and also predict the type of this oxide?

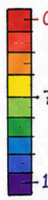
Ans.

$$ZnO + 2HCl \longrightarrow ZnCl_2 + H_2O$$
$$ZnO + 2NaOH \longrightarrow Na_2ZnO_2 + H_2O$$

ZnO is amphoteric in nature because it acts as acid as well as base.

(b) Why $AlCl_3$ is an acidic halide, but NaCl not?

Ans.
- $AlCl_3$ is acidic halide because when it dissolves in water it produces strong acid and causes a decrease in pH.
- NaCl is a neutral halide when it dissolves in water it doesn't change the pH of water.



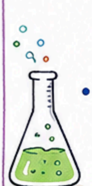
(c) Predict whether the chlorides PCl_5, NCl_3 would be acidic or basic, give reason.

Ans.
- PCl_5 is acidic in nature because it reacts with water and produce HCl and H_3PO_4. Showing acidic property.

$$PCl_5 + 4H_2O \longrightarrow H_3PO_4 + 5HCl$$

- NCl_3 is weakly basic in nature because on hydrolysis it yields NH_3 and HOCl. Since HOCl is a weak acid and NH_3 is relatively strong base so it is base in nature.

$$NCl_3 + 3H_2O \rightarrow NH_3 + 3HOCl$$



(d) Would SO_2 and P_2O_5 react with HCl and H_2SO_4 or with NaOH?

Ans.
- SO_2 has no reaction with HCl and H_2SO_4. It reacts with NaOH and thus, it is acidic in nature.

$$SO_2 + 2NaOH \longrightarrow Na_2SO_3 + H_2O$$

- P_2O_5 reacts as oxidizing agent when reacts with HCl and as dehydrating agent when reacts with H_2SO_4.

$$P_2O_5 + 12HCl \longrightarrow 2POCl_3 + 6H_2O \quad \text{(Oxidizing agent)}$$
$$P_2O_5 + 3H_3SO_4 \longrightarrow 2H_3PO_4 + 3SO_3 \quad \text{(Dehydrating agent)}$$

TEXT BOOK EXERCISE

MULTIPLE CHOICE QUESTIONS

Q.1 Choose the correct one from the given options.

(i) Which scientist first time observed the periodicity in the elements?

(A) J. Newlands (B) L. Meyer
(C) J.W. Döbereiner (D) D. Mendeleev

Answer:
(A) J. Newlands

(ii) Recognize the element if it has 3 electron shells, belongs to "s" block and has 2 electrons in its outer most shell.

(A) Calcium (B) Sodium
(C) Magnesium (D) Potassium

Answer:
(A) Calcium

(iii) Which one do you think is correct about metallic character?

(A) It decreases from top to bottom in a group.
(B) It increases from top to bottom in a group.
(C) It remains constant from left to right in a period.
(D) It increases from left to right in a period.

Answer:
(B) It increases from top to bottom in a group.

(iv) Which property increases as you go down a group in the periodic table?

(A) Atomic radius (B) Electron Affinity
(C) Electronegativity (D) Ionization energy

Answer:
(A) Atomic radius

(v) Which set of the following conditions results in higher ionization energy?

(A) Smaller atom and greater nuclear charge.
(B) Smaller atom and smaller nuclear charge.
(C) larger atom and greater nuclear charge.
(D) larger atom and the smaller nuclear charge.

Answer:
(A) Smaller atom and greater nuclear charge.

Note: The answers are based on fundamental concepts of the Periodic Table.

SHORT QUESTIONS

Q.2 Answer the following short questions.

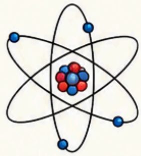
(a) What is 1st ionization energy? Give an example.

Ans. First ionization energy

"The energy needed to remove one electron from each atom in one mole of atoms of the element in the gaseous state to form one mole of gaseous 1^+ ions is known as 1st ionization energy (ΔH_{i1})."

$$Na(g) \rightarrow Na^+(g) + e^- \qquad \Delta H_{i1} = 494\ kJ\ mol^{-1}$$

$$Ca(g) \rightarrow Ca^+(g) + e^- \qquad \Delta H_{i1} = 590\ kJmol^{-1}$$



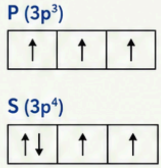
(b) Explain why sulfur has a lower first ionization energy than phosphorus.

Ans. Lower 1st ionization energy of S than P

$_{15}P = 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

$_{16}S = 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$

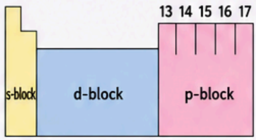
In case of phosphorous 3p subshell is half filled due to stability of half-filled orbitals removal of electron is difficult and I.E is greater. In case of sulphur 3p subshell contain four electrons so, Sulphur's less stable electron configuration compared to phosphorous reduces its I.E.



(c) Why the elements in Group 13 to 17 are called p-block elements?

Ans. p-block elements

Elements in the periodic table can be classified based on the subshells containing their valence electrons the element in group 13 to 17 are called p-block elements because the valance electrons of these elements are in the "p" subshells.



(d) What are the factors that affect electronegativity?

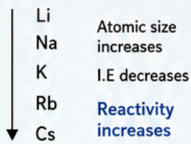
Ans. Factors affecting electronegativity

i. Atomic size
ii. Effective nuclear charge
iii. Shielding effect

(e) What factors are responsible for the the increasing reactivity of alkali metals as you move down the group?

Ans. Factors responsible for increasing reactivity

Moving down the group size of alkali metals increase and tendency of removal of electron increase due to their decrease in I.E that's why reactivity of alkali metals increases down the group.



(f) Why some of the elements show variable oxidation numbers while others do not?

Ans. Reason for variable oxidation numbers

Some element shows variable oxidation number due to d/f-orbital participation, inert pair effect and expanded octet.

(g) Identify the element which is in period 5 and group 15?

Ans. Sb (Antimony)

(h) Why oxides of sodium and magnesium are more ionic than the oxides of nitrogen and phosphorous?

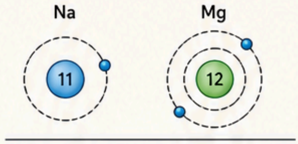
Ans. Reason for ionic nature

Oxides of sodium and magnesium are more ionic than the oxides of nitrogen and phosphorous because sodium and magnesium have large size and law I.E and more electropositive nature and have low electronegativity than nitrogen and phosphorous.

(i) **Give reason for the different chemical reactivities of Na and Mg toward oxygen and chlorine.**

Ans. Reason for different reactivity

Sodium is more reactive than Mg due to its lower ionization energy and more electropositive character and has greater tendency to lose its valance electron than Mg. sodium has single valence electron but Mg has two electrons it must lose two electrons requiring more energy.



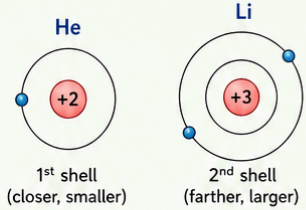
Na	Mg
Lower I.E	Higher I.E
More electropositive	Less electropositive
More reactive	Less reactive

(j) **Why the ionization energy of lithium is much lower than that of helium despite the fact that the nuclear charge of lithium is +3 and that of helium is +2.**

Ans. Reason for lower ionization of energy

$_2He = 1s^2$, $\quad _3Li = 1s^2, 2s^1$

I.E of lithium is much lower than that helium because size of lithium is greater than helium removal of electron from 2nd shell is much easier as compared to 1st shell the outer electron of lithium is farther.



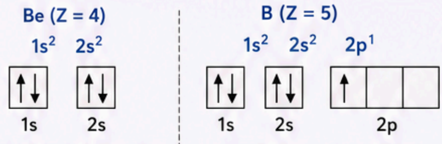
Electron in Li is farther from the nucleus, so it is easier to remove. Hence, I.E of Li is much lower than He.

(k) **The ionization energy of Be (atomic no. 4) is higher than that of B (atomic no. 5), despite the fact that the nuclear charge of Be is +4 and that of B is +5.**

Ans. Reason for higher ionization of energy

$Be = 1s^2, 2s^2$ $\qquad B = 1s^2, 2s^2, 2p^1$

"Be" has higher I.E than "B" because its $2s^2$ electron is more tightly bound (stable full subshell) in other "B" has $2p^1$ electron which is more shielded and feel less nuclear attraction and easily removed.



$2s^2$: stable full subshell (more tightly bound) → higher I.E

$2p^1$: more shielded, less nuclear attraction → lower I.E

(l) **What is common in Na^+, Mg^{2+}, Al^{3+}, Ne^0 and F^-? Arrange them in increasing order of sizes.**

Ans. Common thing

Ten electrons are common in all given species.

Increasing order of sizes

$F^- > Ne^0 > Na^+ > Mg^{2+} > Al^{3+}$

Species	No. of electrons	Nuclear charge (+Z)
F^-	10	9
Ne^0	10	10
Na^+	10	11
Mg^{2+}	10	12
Al^{3+}	10	13

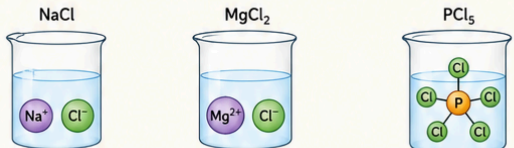
As nuclear charge increases, size decreases.

(m) **Consider the chlorides of sodium, magnesium, and phosphorus: NaCl, $MgCl_2$, and PCl_5**

(i) Classify each of these chlorides as acidic, basic, or neutral.

Ans. Nature of chlorides

NaCl	Neutral chloride
$MgCl_2$	Neutral chloride
PCl_5	Acidic chloride



NaCl: No hydrolysis (neutral)

$MgCl_2$: No hydrolysis (neutral)

$PCl_5 + H_2O \rightarrow POCl_3 + 2HCl$

(forms H^+ in water) → acidic

(ii) For each chloride, briefly explain the reason for your classification, referring to their behaviour when dissolved in water.

Ans. (For answer to this question, see article 1.9.2).