1 PERIODIC TABLE AND PERIODIC PROPERTIES

Student Learning Outcomes [C-11-B-01 to C-11-B-18]

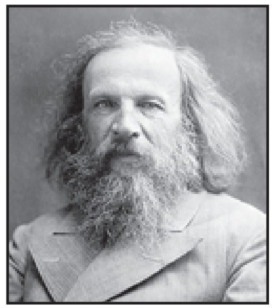
After studying this chapter, students will be able to:

- Explain the arrangement of elements in the periodic table. **(Understanding)**
- Identify the positions of metals, nonmetals and metalloids in the periodic table. **(Understanding)**
- Explain that the periodic table is arranged into four blocks associated with the four sublevels s, p, d, and f. **(Understanding)**
- Recognize that the period number (n) is the outer energy level that is occupied by electrons. **(Understanding)**
- Deduce the electron configuration of an atom from the element's position on the periodic table, and vice versa (based on s, p, d and f subshells). **(Understanding)**
- State that the number of the principal energy level and the number of the valence electrons in an atom can be deduced from its position on the periodic table. **(Knowledge)**
- Deduce the nature, possible position in the Periodic Table and identity of unknown elements from given information about physical and chemical properties. **(Understanding)**
- Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity. **(Application)**
- Explain that vertical and horizontal trends in the periodic table exist for atomic radius, ionic radius, ionization energy, electron affinity and electronegativity. **(Understanding)**
- Explain the trends in the ionization energies and electron affinities of the Group 1 and Group 17 elements. **(Understanding)**
- Recognize that trends in metallic and non- metallic behavior are due to the trends in valence electrons. **(Understanding)**
- Suggest the types of chemical bonding present in the chlorides and oxides from observations of their physical and chemical properties. **(Understanding)**
- Describe (including writing equations for) the reactions, if any, of the oxides (acidic and basic) with water (including the likely pHs of the solutions obtained). **(Understanding)**
- Explain with the help of equations for, the acid / base behavior of the oxides and the hydroxides NaOH, $Mg(OH)_2$ including, where relevant, amphoteric behavior in reactions with acids and bases (sodium hydroxide only) **(Understanding)**
- Explain with equations for, the reactions of the chlorides with water including the likely pHs of the solutions obtained. **(Understanding)**

- Explain the variation in the oxidation number of the oxides and chlorides ($NaCI$, $MgCI_2$ in terms of their outer shell (valence shell) electrons. **(Understanding)**
- Write equations for the reactions of Na and Mg with oxygen, chlorine and water. **(Application)**
- Explain the variations and trends in terms of bonding and electronegativity. **(Understanding)**

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.

1.1 HISTORICAL BACKGROUND



Dmitri Mendeleev arranged elements according to their atomic masses and his table was the first most notable effort in the classification of elements

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 18th century. **Antoine Lavoisier** attempted to classify known elements as metals and nonmetals. 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 of such triads include lithium, sodium, and potassium (^{7}Li, ^{23}Na, ^{39}K).

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 element resembled the first element in properties.

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

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.

In 1913, **Moseley** determined the exact atomic numbers of known elements using X-ray emission, resolving flaws and discrepencies 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.

1.2 MODERN PERIODIC TABLE - MAIN FEATURES

The classification of elements in the modern periodic table helps in the easier understanding of their properties. Following are some of the main features of the modern periodic table:

- Presently, 118 elements are grouped in the table in ascending order of their respective atomic numbers.
- There are seven horizontal rows called **periods** and eighteen vertical columns called **groups**. (In older versions of the table, there were 8 vertical groups were divided into two types of groups: eight A-Groups and ten B-Groups.
- 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.
- Elements in a period show a gradual change in properties moving from left to right in periods.

Other than groups and periods in the periodic table, there are different ways of grouping the elements into various blocks, families and categories just to enhance understanding.

1.3 METALS, NON-METALS AND METALLOIDS

Elements can be broadly classified as metals, nonmetals and metalloids. Metals are elements which tend to lose electrons to form positive ions. Examples are iron, copper, gold and silver. On the other hand, non-metals are elements which tend to gain electrons to form negative ions. The examples are chlorine, sulfur and phosphorous. The metalloids separate the metals and nonmetals in the periodic table. The metalloids exhibit some properties of metals and some of non-metals. Mostly periodic tables have a “stair-step line” on the table identifying the element groups. The line begins at boron (B) and extends down to polonium (Po) including Si, Ge, As, Sb and Te. Elements to the left of the line are considered metals. Elements just to the right of the line exhibit properties of both metals and nonmetals and are termed as metalloids or semimetals. Elements to the far right of the periodic table are nonmetals. The exception is hydrogen, the first element on the periodic table.

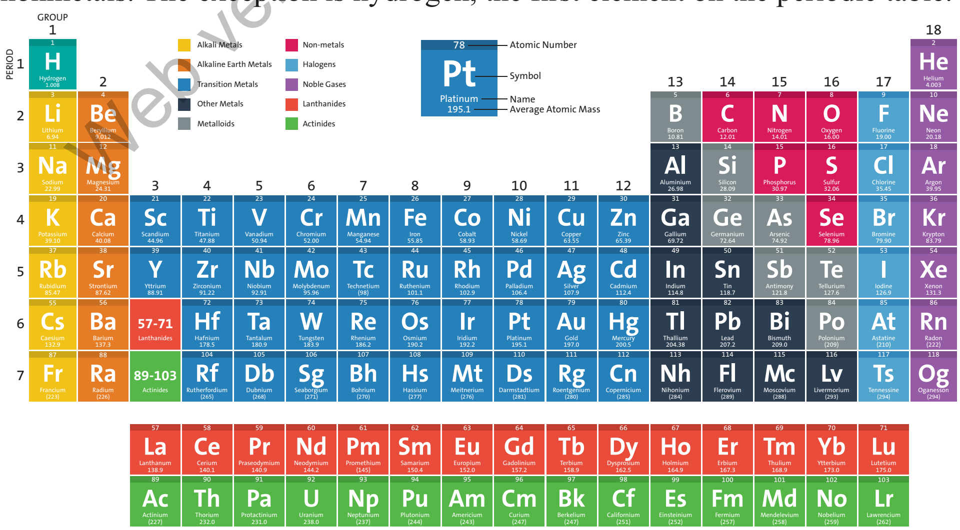


Figure 1.1 Modern periodic table

1.4 BLOCKS IN PERIODIC TABLE

Elements in the periodic table can be classified based on the subshells containing their valence electrons. For instance, the valence electrons of elements in the first two groups are in the **"s"** subshells, placing these elements in the s-block.

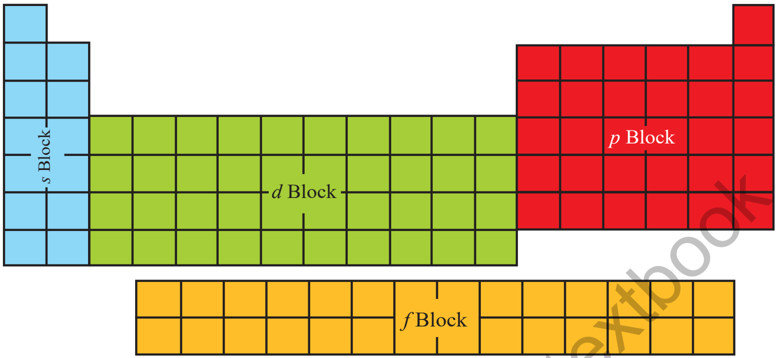


Figure 1.2 Blocks in periodic table

Similarly, transition elements belong to the d-block, and the elements in the two series at the bottom of the table (known as Lanthanides and Actinides) are categorized as f-block elements. The remaining elements in groups 13 to 18, including the inert gases in the last group, belong to the p-block. Knowing the block to which an element belongs provides valuable information about its characteristics, chemical reactivity, oxidation states and other properties such as electronegativity and ionization energy,electron filling, etc.

1.5 FAMILIES IN PERIODIC TABLE

Elements may be categorized according to element families. An element family is a set of elements sharing common properties. There are five famous families of elements in the periodic table:

i) Alkali metals

Elements in the group 1 of the periodic table (Li, Na, K, Rb, Cs, Fr) are known as alkali metals because they produce alkalis when they react with water. Alkali metals have one valence electron. These are the most reactive metals.

ii) Alkaline earth metals

Group 2 elements (Be, Mg, Ca, Sr, Ba, Ra) are metals 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 metals than alkali metals.

iii) Transition metals

The transition metals make up the largest family of elements in the middle of periodic table. They include four series of d-block elements, as well as the lanthanides and actinides (f-block elements) found in the two rows below. Examples include Cr, Fe, Co, Ni, Cu and Zn. They exhibit variable oxidation states. They mostly form coloured compounds.

iv) Chalcogens

The group 16 elements (O, S, Se, Te, Po, Lv) are called Chalcogens because most ores of copper (Greek *chalkos*) are oxides or sulfides. In this group, oxygen & sulfur are non-metals, Se, Te, Po are metalloid and livermorium is a metal.

v) Halogens

Elements in group 17 (F, Cl, Br, I, At, Ts), known as halogens. The term "halogen" means "salt-former" because these elements easily react with alkali metals and alkaline earth metals to form salts. Halogens are highly reactive nonmetals with high electron affinities. Halogens can easily accept one electron to complete their outermost shell.

vi) Noble Gases

The noble gases are a group of unreactive elements present at the extreme right of the periodic table in Group 18 (He, Ne, Ar, Kr, Xe, Rn, Og). Due to their complete outermost shells, they are almost entirely unreactive under normal conditions and rarely form compounds with other elements.

Keep in Mind

Although, noble gases are unreactive, however they have some compounds. An example is compounds of xenon such as xenon hexafluoroplatinate ($XePtF_6$), the word inert gases was changed to noble gases.

Quick Check 1.1

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

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

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.

- The period number indicates shell number (n), representing the number of electron shells surrounding the nucleus. For example, an element **X** in the 3rd period has three electron shells, with its valence electrons located in the 3rd shell. The specific subshell where the valence electrons are found, depends on the element's block. If an element **X** in the 3rd period is in the s-block, its valence electrons are in the 3s subshell.
- The group number indicates the number of valence electrons; for instance, an element **X** in the 3rd period and group 2 has two valence electrons in its outermost shell. Thus, the element **X** in the 3rd period and group 2 (s-block) has two valence electrons in the 3s subshell, which means that **X** would be magnesium (Mg).

Consider another example:

X belongs to group 13 and period 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 3rd shell. The configuration will be:

$$\underset{\text{1st shell}}{\underset{\downarrow}{1s^2}} \qquad \underset{\text{2nd shell}}{\underset{\downarrow}{2s^2\ 2p^6}} \qquad \underset{\text{3rd shell}}{\underset{\downarrow}{3s^2\ 3p^1}}$$

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.

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.

b) Predict the electronic configuration of an element that is in period 4 and group 17 without consulting the periodic table.

1.7 PERIODICITY OF PROPERTIES

The Modern Periodic Law states: "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.

1.7.1 Atomic Radius

The atomic radius is a measure of the size of an atom. It is half of the distance between two identical atoms bonded together. The atomic radius can vary depending on the type of bond (covalent, metallic & vander Waals forces) and the state of the atom. For example, the radius can be different in a covalent bond compared to an ionic bond. The atomic radius is typically measured in picometer (pm) or Angstrom (Å).

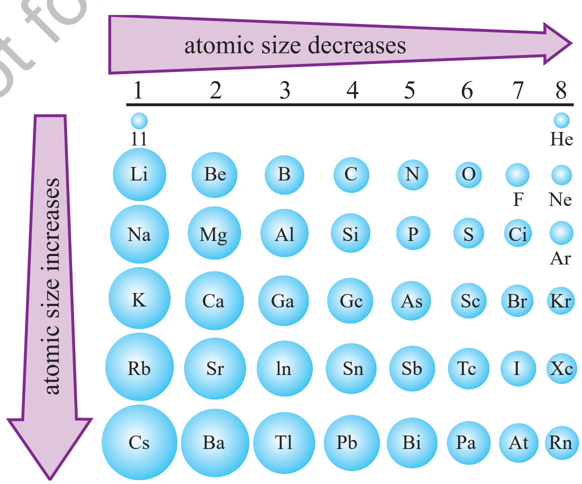


Figure 1.3 Variation in atomic radius across periods and down the groups

Periodic trends in atomic radius:

The factors affecting the atomic radius are: number of shells, effective nuclear charge and shielding effect of inner electrons. 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. Conversely, 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) .

1.7.2 Ionic Radius

The ionic radius is a measure of the size of an ion in a crystal lattice. It is defined as the distance from the nucleus of an ion to the outermost electron shell, measured in picometer (pm) or angstrom (Å). When an atom loses one or more electrons to become a positive ion, it generally becomes smaller than the neutral atom. This is because of the loss of an electronic shell. The nucleus pulls the remaining electrons more strongly. Contrarily when an atom gains one or more electrons to become an anion, it generally becomes larger than the neutral atom. This is because the addition of electrons increases electronic repulsion, as a result the electron cloud expands.

While moving **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.

On the other hand, both cations and anions increase in size as we **move down a group**. This is because the shell 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.

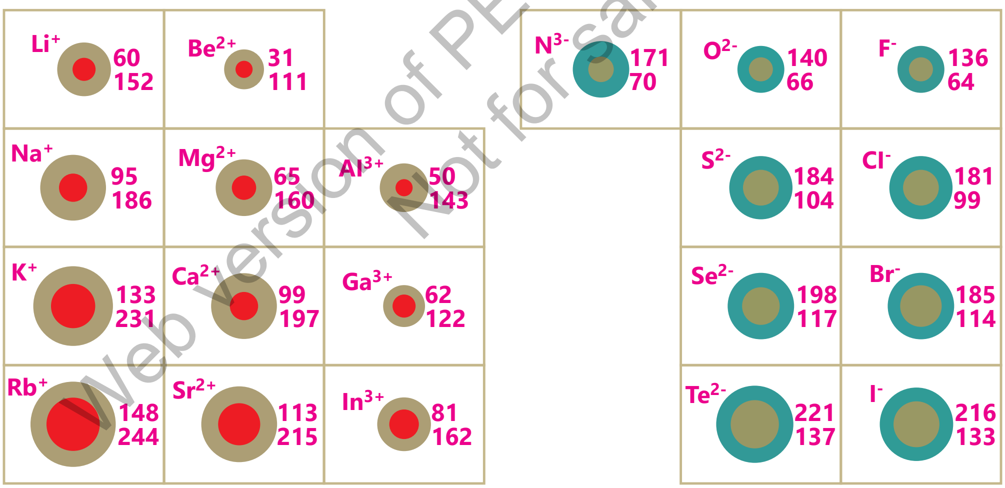


Figure 1.4 Variation in Ionic Radius (Ions are coloured red (cation) and green(anion); parent atoms brown

Quick Check 1.3

a) Which factors affect 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^-.

1.7.3 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})."

$$Na_{(g)} \longrightarrow Na^{+}_{(g)} + e^{-} \quad \Delta H_{i1} = 494 \text{ kJ mol}^{-1}$$

$$Ca_{(g)} \longrightarrow Ca^{+}_{(g)} + e^{-} \quad \Delta H_{i1} = 590 \text{ kJ mol}^{-1}$$

If a second electron is removed from each ion in a mole of gaseous 1+ ions, we call it the 2nd ionization energy, ΔHi2. Again, using calcium as an example:

$$Ca^{+}_{(g)} \longrightarrow Ca^{2+}_{(g)} + e^{-} \qquad \Delta H_{i2} = 1150 \text{ kJ mol}^{-1}.$$

Removal of a third electron from each ion in a mole of gaseous 2+ ions is corresponds to the 3rd ionization energy. Again, using calcium as an example:

3rd ionization energy:

$$Ca^{2+}_{(g)} \longrightarrow Ca^{3+}_{(g)} + e^{-} \quad \Delta H_{i3} = 4940 \text{ kJ mol}^{-1}$$

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.

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.

iii) Electronic arrangement

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

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

b) Oxygen has lower ionization energy than nitrogen due to the half-filled subshell of nitrogen.

$_7N = 1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ (more stable) $\quad \Delta H_{i1} = 1403 \text{ kJ mol}^{-1}$

$_8O = 1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$ (less stable) $\quad \Delta H_{i1} = 1365 \text{ kJ mol}^{-1}$

iv) 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 shielding effect, therefore, lower is the 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. 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.

$_{24}Cr$ = [Ar] $3d^5 4s^1$ (more stable)

$_{25}Mn$ = [Ar] $3d^5 4s^2$ (less stable)

Periodic trends in ionization energy

Going **down in a group**, ionization energy decreases from top to bottom due to increase in the atomic size and shielding effect. In Group I, the ionization energies decrease in the following order: Li > Na > K > Rb > Cs.

In moving from left to right **across a period**, number of shells remains unchanged while the effective nuclear charge increases. Therefore, the ionization energy increases from left to right.

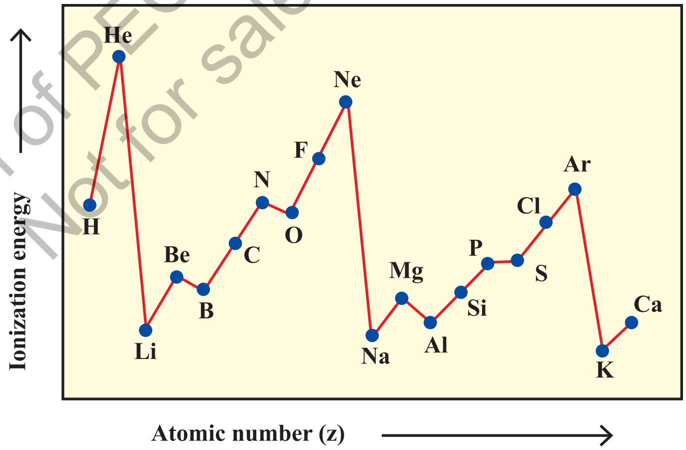


Figure 1.5 Variation in Ionization Energies across periods

The trend of ionization energies of period (1-3) is shown in **Figure 1.5**. It shows that noble gases have the highest values of ionization energy due to complete outermost shell in them, the removal of electron is extremely difficult, whereas alkali metals have lowest values of ionization energy.

Quick Check 1.4

a) Explain with reasoning the following facts about ionization energy:
  i. 1^{st} ionization energy of boron is lesser than beryllium.
  ii. 1^{st} ionization energy of aluminum is lower than magnesium.

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

1.7.4 Electron Affinity (ΔH°_{ea})

The first electron affinity, (ΔH°_{ea1}), 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.

For example:

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \qquad \Delta H^\circ_{ea1} = -348.8 \text{ kJ mol}^{-1}$$

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.

The second electron affinity, ΔH°_{ea2} 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. For example, when first electron is added to a neutral oxygen atom, 141 kJ mol^{-1} energy is released.

$$O_{(g)} + e^- \longrightarrow O^-_{(g)} \qquad \Delta H^\circ_{ea1} = -142 \text{ kJ mol}^{-1}$$

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

$$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} \qquad \Delta H^\circ_{ea2} = +844 \text{ kJ mol}^{-1}$$

It is because the negatively charged ion repels the incoming electron.

Factors affecting electron affinity

Important factors affecting the magnitude of electron affinity 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. Electron affinity values of 'N' and 'P' (group-15), atoms are very low. This is because of the presence of half-filled 'np' subshell 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.

Periodic trends in electron affinity

As the atomic size increases **down the group**, the larger atomic size causes the incoming electron to experience less attraction from the nucleus. Consequently, electron affinity generally decreases down the group. This trend is observed in halogens (At < I < Br < F < Cl). Generally, electron affinities become more negative as we move from left to right in a **period**. This is firstly due to increase in the nuclear charge, which attracts additional electrons more strongly and secondly due to decreasing atomic radius.

Table 1.1 Electron Affinities (kJ/Mol) for Group 1 and Group 17

ELEMENT	ELECTRON AFFINITY (kJ/Mol)	ELEMENT	ELECTRON AFFINITY (kJ/Mol)
Fluorine	-328.0	Lithium	-60.0
Chlorine	-349.0	Sodium	-53.0
Bromine	-324.0	Potassium	-48.0
Iodine	-295.0	Rubidium	-47.0
Astatine	-270.1	Cesium	-46.0

Quick Check 1.5

Explain with reasoning following facts about electron affinity:

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

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

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

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

1.7.5 Electronegativity

Electronegativity is the power of an atom to attract shared pair of electrons toward itself in a molecule. Linus Pauling, 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).

Did You Know?

Linus Pauling is the only person to have received two unshared Nobel Prizes, one for chemistry in 1954 for his work on the nature of chemical bond and one for peace in 1962 for his opposition to weapons of mass destruction.

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.

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. The electronegativity of Li in period 2 is 1.0 and F has a value of 4.0.

Periodic trends in electronegativity

When we move from left to right **along the period**, the electronegativity increases. This is due to increasing nuclear charge and decreasing size. 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 periodic table in **Figure 1.6**.

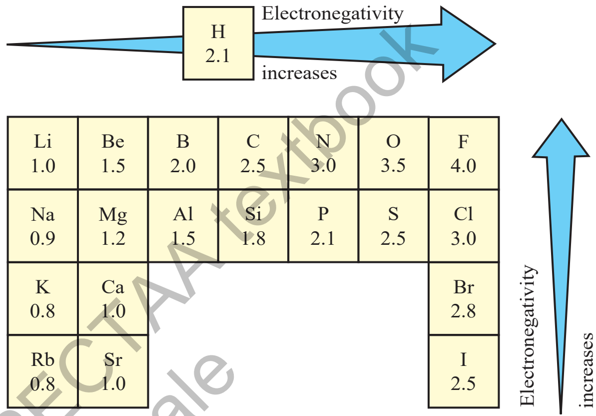


Figure 1.6 Variation of electronegativity in groups and periods

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.

Figure 1.7 provides a summary of all the variation trends in various physical properties of elements in the periodic table.

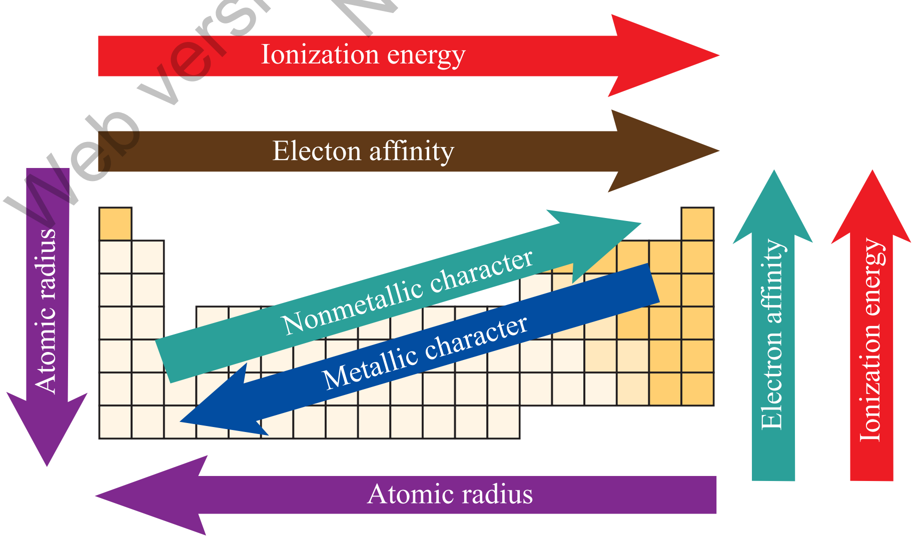


Figure 1.7 Trends in various physical properties in the periodic table

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 noble gas configuration. In contrast, elements on the right side of the table tend to gain electrons. Therefore, elements on the left side of the periodic table are **metals** that form positive ions, while elements on the right side, particularly in the right corner, are **nonmetals** that form negative ions. Hence one can conclude that the metallic character of an element largely depends on its valence shell electronic configuration.

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 across the period, the nuclear charge increases while the atomic size decreases, which results in stronger attraction to the valence electrons making its difficult. Therefore, metallic character decreases from left to right.

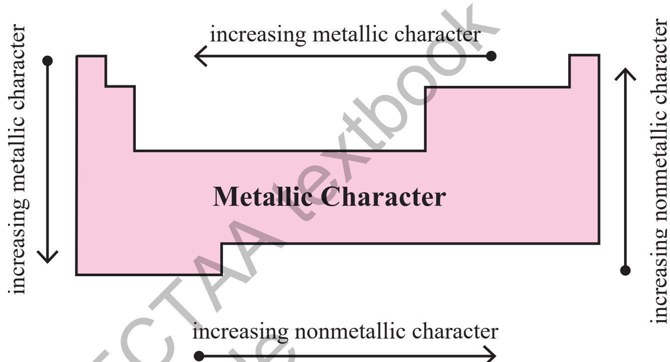


Figure 1.8 General trends for the metallic character of elements

Contrarily metallic character increases as one moves **down in a group** of the periodic table. This is due to the increase in atomic size and the shielding effect. Which reduce the nuclear attraction on the valence electrons. The increase in metallic character (ease of losing electron) makes the element more reactive. Hence Cesium is far more reactive and electropositive than sodium or lithium.

Quick Check 1.6

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

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

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

Magnesium powder burns very rapidly with an intense white flame. This has led to its use in fireworks and S.O.S. flares

$$2Na_{(s)} + 2H_2O_{(\ell)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$
$$Mg_{(s)} + 2H_2O_{(\ell)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$
$$Mg_{(s)} + 2H_2O_{(g)} \longrightarrow MgO_{(s)} + 2H_{2(g)}$$

1.8.2 With Oxygen

Sodium burns in oxygen with a 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.

$$2Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{2(s)}$$

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

$$4Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{(s)}$$

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

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(g)}$$

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

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$
$$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$$

Quick Check 1.7

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

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

1.9 TRENDS IN BONDING IN OXIDES AND CHLORIDES 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. This transition is a result of the increasing electronegativity and decreasing ionic character.

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.

1.9.1 Classification of Oxides

Oxides

Oxides are binary compounds formed by the reaction of oxygen with other elements. The classification of oxides is done into neutral, amphoteric and basic or acidic based on their characteristics.

i) Basic oxides

A **basic oxide** is an oxide that when combined with water gives off an alkali. Metals react with oxygen to give basic oxides. These oxides are usually ionic in nature. Group 1 and 2 form basic oxides when react with oxygen. Examples are: Na_2O, CaO, BaO. Group 2 hydroxides solubility increases down the group so alkalinity also increases down the group.

$$Na_2O_{(s)} + H_2O_{(\ell)} \longrightarrow 2NaOH_{(aq)}$$
$$CaO_{(s)} + H_2O_{(\ell)} \longrightarrow Ca(OH)_{2(aq)}$$

ii) Acidic oxides

An acidic oxide is an oxide that when combined with water gives off an acid. Non-metals react with oxygen to form acidic oxides which are held together by covalent bonds. Silicon dioxide is acidic oxide as it can react with bases. Examples of acidic oxides in period 3 are: P_2O_3, P_2O_5, SO_3, SO_2

$$P_2O_{3(s)} + H_2O_{(\ell)} \longrightarrow H_3PO_{3(aq)}$$
$$SO_{2(g)} + H_2O_{(\ell)} \longrightarrow H_2SO_{3(aq)}$$

Reactions of theses oxides with bases are given below:

$$SiO_{2(s)} + 2NaOH_{(aq)} \longrightarrow Na_2SiO_{3(aq)} + H_2O_{(\ell)}$$
$$P_2O_{3(s)} + 6NaOH_{(aq)} \longrightarrow 2Na_3PO_{3(aq)} + 3H_2O_{(\ell)}$$
$$SO_{2(g)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{3(aq)} + H_2O_{(\ell)}$$

iii) Amphoteric oxides

Amphoteric oxides are oxides that can react with both acids and bases. This means they have the ability to behave as either an acid or a base, depending on the conditions. **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.

$$Al_2O_{3(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(s)} + 3H_2O_{(\ell)}$$
$$Al_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow 2NaAlO_{2(s)} + H_2O_{(\ell)}$$

1.9.2 Classification of Chlorides

Chlorine forms compounds with other elements known as **chlorides.** These chlorides show characteristic behavior when we add them into water, resulting in solutions that can be acidic or neutral.

i) Neutral chlorides

Neutral chlorides are salts that, when dissolved in water, produce a neutral solution with a pH close to 7. At the start of period 3, chloride sodium and magnesium do not react with water. 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. For example,

$$NaCl_{(s)} \longrightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)} \quad (pH=7)$$

$$MgCl_{2(s)} \longrightarrow Mg^{+2}_{(aq)} + 2Cl^{-}_{(aq)} \quad (pH=6.5)$$

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

ii) Acidic chlorides

If we move in period 3, from alumium to sulfur all chlorides react with water to make acidic solution with pH less than 7 this process is called hydrolysis. When $AlCl_3$ is added to water, aluminium and chloride ions in solution. Al^{3+} ion is hydrated and casues a water molecule to lose an H^+ ion, this process is hydrolysis. This turns the solution acidic. The following reaction occurs:

$$AlCl_{3(s)} + 3H_2O_{(\ell)} \longrightarrow Al(OH)_{3(s)} + 3HCl_{(aq)}$$

Other examples of acidic chlorides are given below:

$$SiCl_{4(\ell)} + 2H_2O_{(\ell)} \longrightarrow SiO_{2(s)} + 4HCl_{(aq)}$$

$$PCl_{3(g)} + 3H_2O_{(\ell)} \longrightarrow H_3PO_{3(s)} + 3HCl_{(aq)}$$

Quick Check 1.8

- **a)** ZnO reacts with HCl to give $ZnCl_2$ and with NaOH to give Na_2ZnO_2. Give equations and also predict the type of this oxide?
- **b)** Why $AlCl_3$ is an acidic halide, but NaCl not?
- **c)** Predict whether the chlorides PCl_5, NCl_3 would be acidic or basic, give reason.
- **d)** Would SO_2 and P_2O_5 react with HCl and H_2SO_4 or with NaOH?

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. 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. The oxidation number matches the group number, reflecting the total number of valence electrons. Consider **Table 1.2** for oxidation states of various elements of the periodic table. In the oxides, the oxidation number increases from +1 in Na to +6 in S. In chlorides, the oxidation number increases from +1 in Na to +5 in P. Phosphorus and sulfur exhibit several oxidation numbers because they can expand their octet by exciting electrons into empty 3d orbitals. For instance, in SO_2, sulfur has an oxidation number of +4 because only four electrons are used for bonding, while in SO_3, sulfur has an oxidation number of +6 because all six electrons are used for bonding.

Table 1.2 Oxidation Numbers in Oxides and Chlorides of 3rd Period elements

Oxide	Oxidation Number	Chloride	Oxidation Number
Na in Na_2O	+1	Na in NaCl	+1
Mg in MgO	+2	Mg in $MgCl_2$	+2
Al in Al_2O_3	+3	Al in $AlCl_3$	+3
Si in SiO_2	+4	Si in $SiCl_4$	+4
P in P_4O_{10}/ P in P_4O_6	+5/+3	P in PCl_5	+5
S in SO_3	+6	P in PCl_3	+3
S in SO_2	+4	S in SCl_2	+2

Quick Check 1.9

a) Calculate the oxidation number of sulfur in SO_2 and SO_3.

b) Why some p block elements show variable oxidation state?

EXERCISE

MULTIPLE CHOICE QUESTIONS

Q.1 Four choices are given for each question. Select the correct choice.

I. Which scientist first time observed the periodicity in the elements?

a) J. Newlands b) L. Meyer

c) J.W. Döbereiner d) D. I. Mendeleev

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

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.

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

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

VI. Which of the following atoms show more than one (variable) oxidation states?

a) Sodium
b) Magnesium
c) Aluminum
d) Phosphorous

VII. Which is the correct general trend in the variation of electron affinity in a group?

a) It becomes less negative from top to bottom.

b) It becomes more negative from top to bottom.

c) It remains the same.

d) It has no definite trend and changes irregularly.

VIII. What is the oxidation state of sulfur in the sulfate ion (SO_4^{2-})?

a) + 4
b) + 2
c) + 6
d) 0

IX. Which is the correct trend in variation of electronegativity along a period of the periodic table?

a) It decreases from left to right across a period.

b) It increases from left to right across a period.

c) It remains constant.

d) It has no definite trend.

X. The atomic radius generally…….. across a period in the periodic table.

a) Increases
b) Decreases
c) Remains constant
d) First increases then decreases

XI. Which one of the following elements has the highest ionization energy?

a) Sodium (Na) b) Magnesium (Mg)

c) Aluminium (Al) d) Argon (Ar)

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

a) What is 1^{st} ionization energy? Give an example.

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

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

d) What are the factors that affect electronegativity?

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

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

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

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

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

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?

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.

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

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

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

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

DESCRIPTIVE QUESTIONS

Q.3 Write equations for the reactions of Na and Mg with oxygen, chlorine, and water. Compare the reactivity of both elements with these in terms of metallic character.

Q.4 Explain with the help of equations, acidic and basic behavior of oxides and chlorides.

Q.5 Describe the factors affecting and periodic trends of electron affinity.

Q.6 Define ionization energy. Discuss the factors affecting and periodic trends of ionization energy.