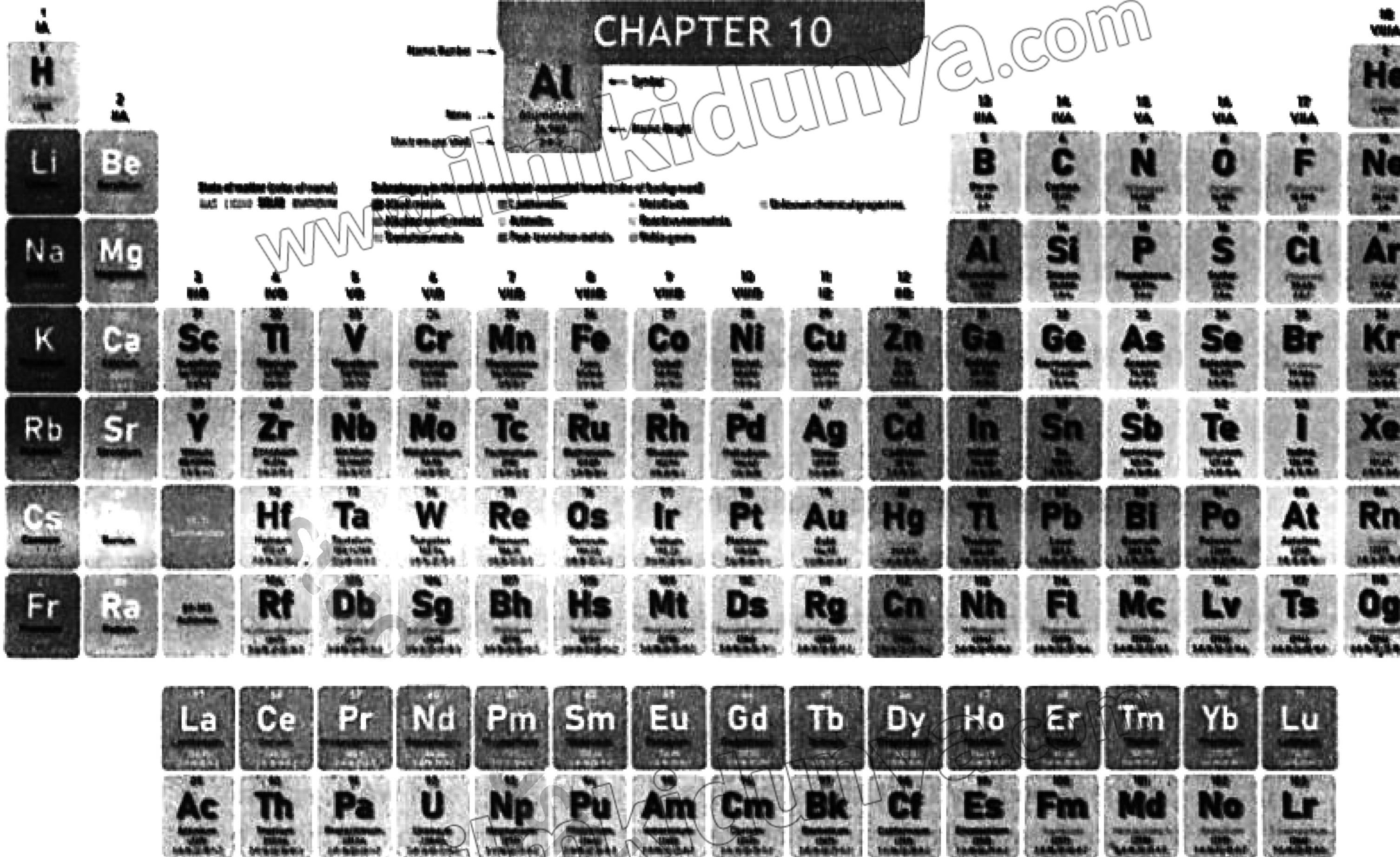
CHAPTER 10



PERIODIC TABLE

SLOs: After completing this lesson, the student will be able to:

1. Explain the arrangement of elements in the periodic table.
2. Explain that the periodic table is arranged into four blocks associated with the four sub-levels- s, p, d, and f.
3. Recognize that period number (n) is the outer energy level that is occupied by electrons.
4. State that the number of the principal energy level and the number of valence electrons in an atom can be deduced from its position on the periodic table.
5. Identify the position of metals, non-metals and metalloids in the periodic table.
6. Explain that vertical and horizontal trends in the periodic table exists for atomic radius, ionic radius, ionization energy, electron affinity and electronegativity.
7. Recognize that trends in metallic and non-metallic behavior is due to the trends in valence electrons.
8. Deduce the electron configuration of an atom from the element`s position in the periodic table, and vice versa (based on s, p, d, and f sub-shells)
9. Write equations for, the reactions of Na and Mg with oxygen, chlorine, and water.
10. Explain the variation in the oxidation number of the oxides and chlorides (NaCl, $MgCl_2$) in terms of their outer shell (valence shell) electrons.
11. 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).
12. Explain with the help of equations for, the acid / base behavior of the oxides and hydroxides NaOH, $Mg(OH)_2$ including amphoteric behavior where relevant, in reactions with acids and bases (NaOH only).
13. Explain with equations for, the reactions of the chlorides with water including the likely pHs of the solution obtained.
14. Explain the variations and trends in terms of bonding and electornegativity.

15. Suggest the type of chemical bonding present in the chlorides and oxides from observation of their physical and chemical properties.
16. Predict their characteristic properties of an element in a given group by using knowledge of chemical periodicity.
17. Deduce the nature, possible position in the periodic table and identity of unknown elements from given information about physical and chemical properties.
18. Explain the trends in the ionization energy and electron affinity of the Group 1 and Group 17 elements.

One of the most important functions is to search for order. This instinct led to the discovery of a systematic arrangement of elements. It is considered the most important achievement in chemistry. When Moseley discovered atomic numbers in 1913, it was discovered that atomic numbers could provide a basis for the systematic arrangement of elements. Thus, the modern periodic table shows the sequence of elements in order of increasing atomic number and their repeating properties. A table that shows the systematic arrangement of the elements is called a periodic table. It is based on the periodic law, which states that when elements are arranged in order of increasing atomic number, their properties repeat periodically

10.1 PERIODS AND GROUPS OF ELEMENTS.

The most commonly used form of the periodic table is shown in Figure 10.1. Based on the periodic law, the elements are arranged in seven horizontal rows called periods. Elements with similar chemical properties are in the same vertical columns. These columns are called groups. Note that the elements are listed in ascending atomic number, from left to right and from top to bottom. Hydrogen (H) is in the upper left corner. Helium (He), atomic number 2, is in the upper right corner. Lithium (Li), atomic number 3, is at the left end of the second row.

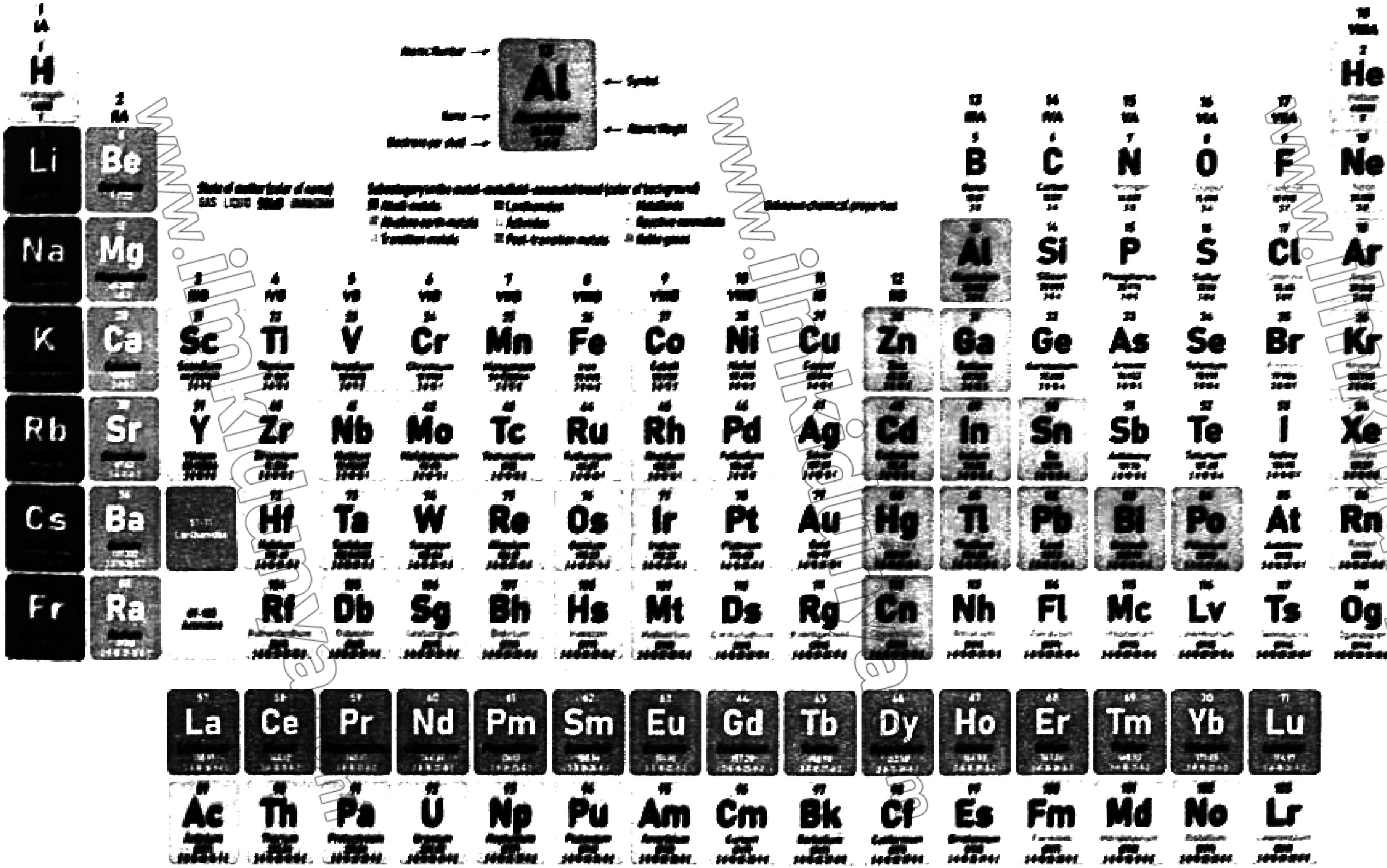
Periods contain a variable number of elements. How many periods can you find in the periodic table? The number of elements per period varies from 2 in period 1 to 32 in period 7. The first three periods are called short periods and the rest are called long periods. The properties of the elements inside a period change gradually as you move from left to right in the loop. However, as you move from one period to the next period, the pattern of features within an period repeats itself. This is in accordance with the periodic law.

Each vertical column of elements in the periodic table is called a group or family. How are groups numbered?

Two numbering systems are often used to designate groups. In the traditional system and the old IUPAC, the letters A and B are used. The first two groups are IA and IIA, while the last six groups are IIIA to VIIIA and the middle groups are in group B. The International Union of Pure and Applied Chemistry (IUPAC) in 1988 decided that the groups would be 1-18 from left to right.

Elements in the same group have the same number of valence electrons. The group number indicates the number of valence electrons in the element. For example, group 1 and group 2 elements have 1 and 2 valence electrons, respectively. Group 13 elements have 3, group 14 elements have 4, group 15 elements have 5 valence electrons, etc. It is important to note that in groups 13-18, the group number is equal to the number of valence electrons plus 10.

Figure 10.1: Periodic Table of Elements



Group A elements are called **normal** or **representative elements**. They are also called main group elements. Group B elements are called **transition elements.**

Groups of elements in the periodic table have been given group names based on the topmost element in the group. Some groups have been given special names. For example, metallic elements in Group 1 are called **alkali metals.** Group 2 elements are called **alkaline earth metals.** Group IIIA or Group 13 elements are called boron family. Group VIA or 16 are called **oxygen family.** They are also called **chalcogens.** The elements in Group 17 or VIIA are **halogens.** The Group 18 or VIIIA elements are called **noble gases** because they do not readily undergo chemical reactions.

10.2 BLOCKS OF ELEMENTS IN THE PERIODIC TABLE

The block of an element refers to the position of the periodic table based on their respective valence sub-shells, which may be s, p, d, or f. Accordingly there are four blocks in the periodic table, s-block, p-block, d-block, and f-block.

s-Block

The elements presents in Groups 1 and 2 (except He) are called s-block elements. They contain their valence electrons in s-sub shells. Alkali and alkaline earth metals are present in s-block.

p-Block

The elements present in Groups 13 to18 are called p-block elements. They contain their valence electrons in p-sub shells. This block contains metals both metals and non-metals.

d-Block

Elements present in Groups 3 to 12(or sub-group B) are called d-block elements. They contain their valence electrons in d-sub shells. These elements are called as transition metals.

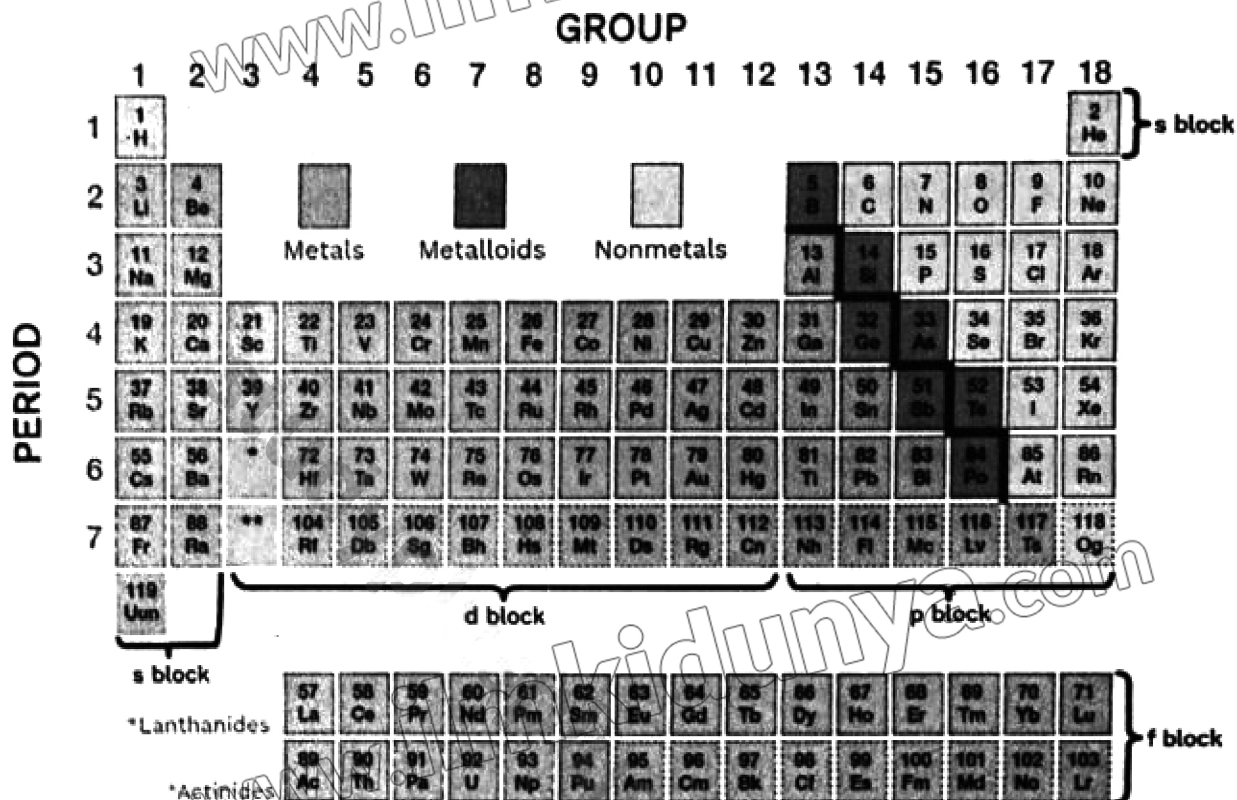
f-Block

Two rows of elements located at the bottom of the periodic table consist of f-block elements. They contain their valence electrons in f-sub shell. These elements are also called lanthanides and actinides or inner transition metals.

10.3 POSITION OF AN ELEMENT IN THE PERIODIC TABLE

The electronic configuration of an element provides basic information about its position in the periodic table. By examining the valence electron configuration, n-value, and sub-shell type, you can identify an element's group, period, and block. To determine the group of an element, you need to know the number of electrons in the valence electrons. For example, if the electron configuration of the valence shell is $2s^2,2p^3$. Since the total number of valence electrons is 2+3= 5, it belongs to the Group VA or 15. If an element belongs to p-block, their group number is equal to the total number of valence electrons plus 10. The n value of the valance shall of an element indicates its horizontal row on the periodic table. The periodic number is equal to the largest n value of the valence shell. For example, if the valence electrons of an element are in the L shell (n=2), then it belongs to the 2nd period. Similarly, if n=3, it belongs to the 3rd period.

Figure 10.2: Periodic Table of Elements



You can determine the position of an element in the periodic table from its valence shell electronic configuration. Period number of an element indicates n value of the valence shell. Whereas total number of electrons in the valence shells represent the group number of s-block elements. But the total number of electrons in the valence shells plus 10 represents the group no of p-block elements.

Characterestic properties of elements in a group

A chemical periodicity refers to the repeating pattern of the properties of elements in the periodic table.

Elements within the same group generally have similar chemical properties because they have similar electronic configurations.

You can predict the properties of elements in a particular group based on their position in the periodic table:

For example:

- Atomic size gradually increases from top to bottom in a group.

  Therefore, elements lower in the group have larger atomic radii than elements higher in the group.

- Ionization energy decreases within a group.

Therefore, elements lower in the group have lower ionization energies than elements higher in the group.

- Electronegativity, decreases down the groups.

  Therefore, elements further down in the group are less electronegative than elements higher up in the group.

- Metallic properties increase as you move down the group.

  Therefore, elements lower in the group have stronger metallic properties than elements higher in the group.

- Chemical reactivity: Elements in the same group generally have similar chemical reactivity because they have similar electron configurations in their valence shells.

  Therefore, elements further down the group exhibit similar chemical behavior to the elements above them in the group.

Example 10.1: Find out the position of the following elements in the periodic table;

(a) Nitrogen (atomic number 7) (b) Oxygen (atomic number 8)

Problem Solving Strategy:

Write the electron configuration of the element. Identify the valence shell configuration. The n value of an s or p sub-shell represents the period number and the total number of electrons in the valence shell is equal to the group number.

Solution:

a) Electronic configuration of N = $1s^2, 2s^2, 2p^3$

   Valence shell has configuration = $2s^2, 2p^3$

   Period number = 2

   Group number = 2 + 3=5

   $\therefore$ Nitrogen is present in the 2nd period of Group V-A

   (5 + 10 = Group 15)

b) Electronic configuration of oxygen = $1s^2, 2s^2, 2p^4$

   Valence shell has configuration = $2s^2, 2p^4$

   So, Period number = 2

   Group number = 2 + 4 = 6

   $\therefore$ Oxygen is present in the 2nd period of Group VI-A

   (6 + 10 = Group 16)

Do You Know?

What Makes Up >99% of the Normal Matter in the Universe?

Helium and hydrogen make up most of the universe. Both of them account for 98% of all matter, being roughly 73% hydrogen, and 25% helium. All the other elements make up the remaining 2% of matter. The next in the list is oxygen, making up for a tiny 0.05%. Other atomic components in this order of magnitude are neon, nitrogen, carbon and silicon.

You can also determine the valence shell electronic configuration of an element from its position in the periodic table. The period number of elements indicates the n value of the valence shell.

Example 10.2: Obtaining the valence shell electronic configuration of

(a) Phosphorous (b) Neon

Problem Solving Strategy:

Remember that

Period number = n value of valence shell

Group number = number of valence electrons

Distribute the electrons in the sub-shells of the valence shell.

Solution:

a) The periodic number of phosphorus is 3, so n = 3, so the valence shell is M, which has 2s and 3p sub-shells. The group number is V, so there are 5 electrons in the valence shell. 2 electrons fill the 3s sub-shell and the remaining 3 fill the 3p sub-shell. Therefore, the electronic configuration of the valence shell is $3s^2$, $3p^3$.

b) The periodic number of neon is 2. So, n = 2 and the valence shell is L. The valence electrons exist in the 2s and 2p sub-shells. The group number of neon is VIII. This means that there are 8 electrons in the valence shell. 2 electrons fill the 2s sub-shell and the remaining 6 fill the 2p sub-shell. So, the electron configuration of the valence shell of neon is $2s^2$, $2p^6$.

Concept Assessment Exercise 10.1

1. Find the valence shell electronic configuration of Mg and Cl from their position in the periodic table.
2. Find the position of K (atomic no.19) and S (atomic no.16) in the periodic table.

10.4 METALS, NON-METALS AND METALLOIDS

The periodic table provides a general framework for organizing elements.

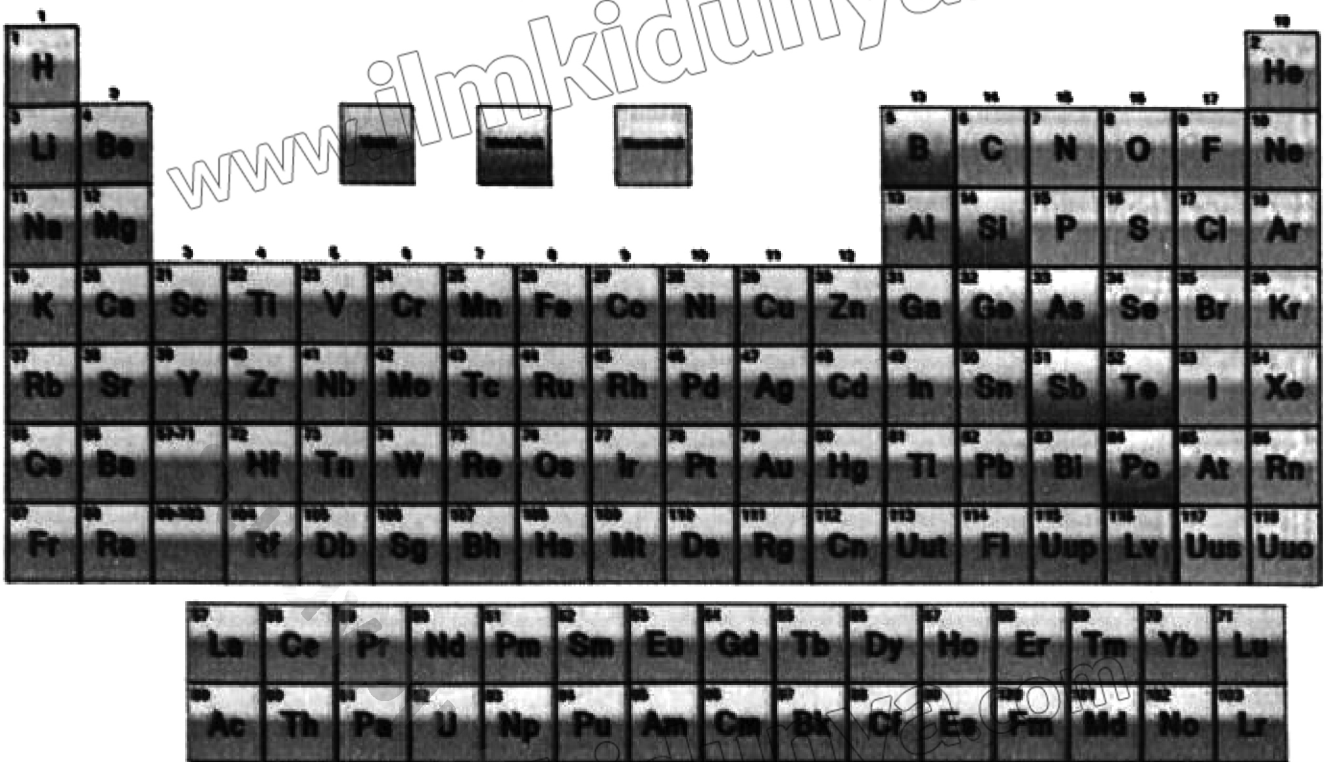
The elements are generally classified in three categories on the bases of their properties and their position in the periodic table 10.3.

1. Metals
2. Non-metals
3. Metalloids

Metals appear on the left hand side of the periodic table, shown by grey colour. Metals are good conductors, malleable and ductile.

Non-metals are present on the right side of the periodic table, shown by yellow colour. Non-metals are bad conductors, non- malleable and non-ductile. A blue leader line separates metals from non- metals. Metals are on left side of the leader line (except hydrogen, which is non-metal), non-metals are on the right side of the line. The

Table 10.3: Position of Metals, Non-metals and Metalloids in the Periodic Table



elements adjacent to this leader line shown by blue colour are metalloids. Metalloids possess properties of both metals and non-metals. They are also called semi-metals.

10.5 PERIODICITY OF PROPERTIES

Remember that the electronic configuration of elements shows a periodic fluctuation with increasing atomic number. Therefore, the physical and chemical properties of the elements vary in a periodic manner. Elements with similar valence shell electronic configurations are placed in the same group, one below the other. Chemical properties depend on the electronic configuration of the valence shell. Because all elements in a given group have a similar valence shell electronic configuration, they have similar chemical properties. Physical properties depend on the size of atoms. Because the sizes of atoms change gradually from top to bottom in a group (see section 2.1.4). Therefore, the elements show a gradation of physical properties within the same group. In the period of the periodic table, the number of electrons in the valence shell increases gradually from left to right. Their chemical and physical properties also differ in the same way.

10.5.1 Electron Affinity

Electron affinity explains the anion formation. **Electron affinity is defined as the amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uni-negative gaseous ion. It is also called first electron affinity.**

$$X_{(g)} + e^- \longrightarrow X^-_{(g)} + \text{electron affinity}$$

Figure 10.7 shows electron affinities of main group elements.

Three factors effect electron affinity.

1. Nuclear charge 2. Atomic size 3. Shielding effect

1. Nuclear charge:

Electron affinity tends to increase from left to right in the periodic table. This is mainly due to the increase in nuclear charge. As the number of protons in the nucleus increases, the attraction to electrons also increases, making it easier for the atom to accept an additional electron.

2. Atomic size:

Electron affinity generally decreases from top to bottom in a group. This is due to the increase in atomic size. Larger atoms have electrons farther from the nucleus and the attraction between the nucleus and the outer electrons is weaker. Therefore, it is more difficult for larger atoms to accept extra electrons.

3. Shielding effect:

The electrons in the inner shells partially shield the outer electrons from the attraction of the nucleus. The more inner shells there are, the more the outer electrons are shielded. This shielding effect can lead to a decrease in electron affinity.

As you move from left to right in a period, electron affinity increases. This is due to an increase in nuclear charge and a decrease in atomic size, which binds the extra electron more tightly to the nucleus. The shielding effect remains the same in a period. The electron affinity decreases from top to bottom in a group. This is due to an increase in the shielding effect. Due to increased shielding effect and increase in atomic size, the added electron binds less tightly to the nucleus. As a result, less energy is relased.

H -73							He 0
Li -60	Be 0	B -27	C -122	N +7	O -141	F -328	Ne 0
Na -53	Mg 0	Al -44	Si -134	P -71.7	S -200	Cl -349	Ar 0
K -48	Ca 0	Ga -29	Ge -120	As -77	Se -195	Br -325	Kr 0
Rb -47	Sr 0	In -29	Sn -121	Sb -101	Te -190	I -295	Xe 0
Cs -45	Ba 0	Tl -30	Pb -110	Bi -110	Po -180	At -270	Rn 0

Figure 10.7 electron affinities of main group elements

10.6 TRENDS IN METALLIC AND NON-METALLIC BEHAVIOURS

The valence electrons play a key role in determining the chemical and physical behavior of elements.

10.6.1 Metallic Behaviour

Metals are usually found on the left side of the periodic table. They tend to easily lose valence electrons and form cations to achieve stable electron configurations. This behavior is due to the following reasons.

Metals usually have 1-3 valence electrons. Their ionization energy is usually low. Therefore, metals can easily lose these electrons. In addition, metals have low electronegativity and do not tend to attract electrons.

10.6.2 Non-Metallic Behaviour

Non-metals are usually found on the right side of the periodic table. Their non-metallic behavior is due to their tendency to gain or share valence electrons to achieve a stable structure. Non-metals usually have 5-7 valence electrons. They need a few more electrons to fill their valence shell. Non-metals have high ionization energy. This makes them less likely to lose electrons and more likely to share or gain electrons. Non-metals also have high electronegativity, which allows them to easily gain electrons or share electrons to form covalent bonds.

10.6.3 Trends in Groups

Group 1 alkali metals have a strong metallic character. Group 17 (VII A) halogens have a strong non-metallic character. As you move down the group, the metallicity of the group 1 element increases. This is due to the increase in atomic radii and their ability to easily lose electrons. But as we go below group 17, the non-metallic behaviour decreases. This is due to the increase in atomic radii and decrease in their ability to attract electrons. In group 14 (IV A), which consists of carbon (C), silicon (Si), germanium (Ge), tin (Sn), and lead (Pb), the non-metallic character decreases and the metallic character increases down the group. Carbon is a typical non-metal, silicon and germanium are metalloids, tin and lead are typical metals.

10.6.4 Electronegativity and Type of Bond

The electronegativity difference (Δ EN) of the bond between two bonded atoms gives a rough indication of the expected nature of the bond and hence bond type. When the difference is greater than 1.8 the bond is ionic. On the other hand, when the difference is 0.4 to 1.8 the bond is polar. When the difference is less than 0.4 the bond is covalent. This means, the ionic character in a bond increases with the increase in the difference of electronegativities of bonded atom.

The electronegativity difference between Group 1 and 2 metals and groups 16 and 17 non-metals is large, therefore bonding between them is ionic. For example in NaCl the Δ EN = 3.16-0.93 = 2.23, so the bond in NaCl is ionic.

In MgO the the Δ EN = 3.44-1.31 = 2.13, so the bond in MgO is ionic.

In C-H the Δ EN = 2.55-2.2 = 0.35, so the C-H bond is non-polar covalent.

In H-F the Δ EN = 3.98-2.2 = 1.78, so the H-F bond is polar covalent.

10.6.5 Trends in Chemical Properties

We will discuss some reactions of typical metals of Group 1 and Group 2, such as sodium and magnesium.

a) Reaction with oxygen:

Sodium is a silvery white soft metal. It is an extremely reactive element. It readily combines with oxygen, chlorine, and water. In a limited supply of oxygen, sodium burns to form sodium oxide(Na_2O). But in excess of oxygen, it forms pale yellow sodium peroxide(Na_2O_2)

$$2Na + O_2 \rightarrow Na_2O$$
$$2Na + O_2 \rightarrow Na_2O_2$$

Magnesium is relatively hard and is also an extremely reactive element. When ignited, Mg burns in oxygen with an intense white flame to give solid magnesium oxide (MgO).

$$2Mg + O_2 \rightarrow 2MgO$$

b) Reaction with chlorine

Sodium burns in chlorine with a bright yellow flame, producing a white solid, sodium chloride.

$$2Na + Cl_2 \rightarrow 2NaCl$$

Magnesium also burns in chlorine with intense white flame, producing a white solid, magnesium chloride.

$$Mg + Cl_2 \rightarrow MgCl_2$$

c) Reaction with water

Sodium reacts violently with cold producing hydrogen. The reaction occurs with a light explosion.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Magnesium slowly reacts with cold water to form a thin layer of magnesium hydroxide on its surface, which forms a protective coating that stops further reaction. In steam, however, magnesium burns with a white flame, forming magnesium oxide and hydrogen.

$$Mg + 2H_2O \rightarrow MgO + H_2$$

10.6.6 Variation in Oxidation Number of Oxides and Chlorides

Oxidation number of an element in oxides and chlorides corresponds to the number of electrons used for bonding. The maximum oxidation number of an element is the same as the total number of valence electrons. It is always positive, since oxygen and chlorine are more electronegative than other elements. For example, in period 3, which consists of Na, Mg, Al, Si, P,S, and Cl. In the oxides of these elements, Ma_2O, MgO, Al_2O_3, SiO_2, P_2O_5, SO_3 and Cl_2O_7, the maximum oxidation number increases in the period from +1 in Na, +2 in Mg, +3 in Al,+4 in Si, +5 in P, +6 in S and +7 in Cl.

Phosphorous and sulphur show variable oxidation number because they can expand their octet e.g, in SO_2, S has oxidation number +4 , but in SO_3 , S has oxidation number +6

Reactions of oxides with water

Metal oxides are generally basic in character. This is because they produce bases in water (pH>7). Non-metallic oxides are generally acidic since they produce acids in water (pH< 7).

$$Na_2O + H_2O \rightarrow NaOH$$
$$MgO + H_2O \rightarrow Mg(OH)_2$$
$$CO_2 + H_2O \rightarrow H_2CO_3$$
$$P_4O_{10} + 6\,H_2O \rightarrow 4H_3PO_4$$

The basic character of oxides decreases as we move along a period. This indicates decrease in metallic character. Also the elecronegtivity difference between these elements and oxygen decreases as we move across the period. So, the bonding character also changes from ionic to covalent. Oxides of sodium and magnesium are ionic compounds. They possess high melting and boiling points. They are soluble in water. Their aqueous solutions are good conductor of electricity. Other oxides are covalent in nature. Oxides of P, S and Cl exhist as gaseous molecules. They have low melting and boiling points.The elements of the third period form oxides such as Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10}, SO_2, Cl_2O_7. These oxides turn from strongly basic through weakly basic, amphoteric, weakly acidic to strongly acidic. See figure 10.7 to understand this trend.

Group	IA	IIA	IIIA	IVA	VA	VIA	VIIA
3rd period	Na	Mg	Al	Si	P	S	Cl
Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_2	Cl_2O_7
Nature of oxide	Strongly basic	Basic	Amphotric (both acidic and basic)	Weakly acidic	Acidic	Strongly Acidic	Very strongly acidic

Table 10.7: Nature of oxides of elements in the period No. 3

Oxide and hydroxide of sodium are strongly basic. They react with acids to from salt and water.

$$Na_2O + HCl \rightarrow NaCl + H_2O$$
$$NaOH + HCl \rightarrow NaCl + H_2O$$

Similarly, oxide and hydroxide of magnesium are also strongly basic, but less basic than those of sodium.

$$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$$
$$Mg(OH)_2 + H_2SO_4 \rightarrow MgSO_4 + 2\,H_2O$$

Oxide and hydroxide of aluminium are amphoteric. They react with both acids and bases

$$Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2MaAl(OH)_4 \text{ (sodium aluminate)}$$
$$Al_2O_3 + 6HCl \rightarrow 3AlCl_3 + 3H_2O$$

10.6.7 Reactions of Chlorides with water

Metal chlorides, such as sodium chloride and magnesium chloride are ionic in nature. They dissolve in water forming a neutral solution. These solutions have pH = 7. On the other hand aluminium chloride ($AlCl_3$), silicon tetra chloride ($SiCl_4$) and phosphorous penta chloride (PCl_5) are covalent compounds. They are also soluble in water, but they react more vigorously with water forming acidic solution. These solutions have low pH.

$$AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl \qquad pH = 3$$
$$SiCl_4 + 4H_2O \rightarrow Si(OH)_4 + 4HCl \qquad pH = 0$$
$$PCl_5 + 4H_2O \rightarrow H_3PO_4 + 5HCl \qquad pH = 0$$

10.6.8 Type of chemical bonding in chlorides and oxides

By looking at their physical and chemical properties, we can infer the types of chemical bonds present in compounds. Ionic compounds usually have higher melting points and boiling points, dissolve in water, and conduct electricity when molten. Covalent compounds often have lower freezing and boiling points, are insoluble in water, and do not conduct electricity. Specific properties and reactivity help distinguish between ionic and covalent bonding, and in some cases a combination of both bond types (ionic/covalent nature) may be present.

10.6.9 Trends in ionization energies and electron affinities of Group 1 and Group 17 elements

As you go down the Group 1, the ionization energy decreases. This is due to two factors.

(i) Increase in shielding effect

(ii) Increase in atomic radii

Both factors reduce the attraction between the nucleus and the valence electrons. Therefore, removing electrons requires less energy. As you move down from the 17th group, the electron affinity decreases. This is also due to the same factors described above. These factors reduce the attraction between the nucleus and the incoming electron. Although fluorine is the most electronegative element, its electron affinity is lower than that of chlorine. This is because the atomic size of F is much smaller than that of Cl. Thus, the incoming electron is repelled by the electron of F atom.

10.6.10 Identification of an unkown element

Suppose we have an unknown element with the following properties:

- Atomic Number: 19
- A highly reactive metal that reacts violently with water to produce hydrogen gas.

The atomic number (19) indicates that the unknown element has electronic configuration is $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$, which indicates this element belongs to Group 1 of the periodic table and is present in the 4th period. It is potassium (K). Its position is confirmed by its electronic configuration and its reaction with water. Because alkali metals react violently with water. This element belongs to Group 1 of the periodic table and is therefore an alkali metal.

- Example 2: Consider an unknown element with the following properties: atomic number: 17
- Exists as diatomic molecules in elemental form.
- Reacts with metals to form a salt.

The atomic number (17) indicates that the electronic configuration of unknown element is $1s^2$, $2s^2, 2p^6, 3s^2, 3p^5$ indicates this element belongs to Group 17 of the periodic table and is present in the 3rd period. This element is chlorine (Cl). It exists as diatomic molecule, Cl_2 and reacts with alkali metals to forms salts, e.g., NaCl

Example 10.5: Identifying position of an unknown element in the periodic table.

Suppose you have an unknown element having atomic number 35, and you want to determine its position in the periodic table.

Problem Solving Strategy

1. Write its electronic configuration.
2. Use valence electronic configuration to locate its position i.e., find its group and period.

Solution

Electronic configuration : $1s^2, 2s^2, 2p^6, 3S^2, 3p^6, 4S^2, 3d^{10}, 4p^5$

Valence shell electronic configuration is $4s^2, 4p^5$, which shows it contains 7 electrons in its valence shell. So this unknown element must placed in Group VII A or 17. The n value of valence shell is 4. As n value of valence sub-shell is 4, this unknown element must lie in the 4th period in the periodic table.

Based on its position in the periodic table, you can predict its properties. For example, does it have a higher or lower melting point, density, and reactivity than the element above or below it? You can also predict its chemical behavior. For example, you can predict the following properties:

i) It is a non-metal.
ii) It must be a diatonic molecule.
iii) It forms an acidic oxide which is more acidic than that of chlorine.
iv) Its maximum oxidation state would be 7
v) combines with metals to form salts.
vi) forms acid with hydrogen etc.

Key Points

- Horizontal rows in the periodic table are called periods.
- Vertical columns in the periodic table are called groups or family.
- Group A elements are called normal or representative elements. They are also called main group elements.
- Group B elements are called transition elements.
- Group 1 are called alkali metals.

- Group 2 elements are called alkaline earth metals.
- Group IIIA or Group 13 elements are called boron family.
- Group VIA or 16 are called oxygen family. They are also called chalcogens.
- The elements in Group 17 or VIIA are halogens.
- The Group 18 or VIIIA elements are called noble gases because they do not readily undergo chemical reactions.
- Period number of an element indicates n value of the valence shell.
- The total number of electrons in the valence shells represent the Group number of an element.
- Metals are good conductors, malleable and ductile.
- Non-metals are bad conductors, non- malleable and non-ductile
- The reduction in force of attraction between nucleus and the valence electrons by the electrons present in the inner sub-shells is called shielding effect.
- Ionization energy is defined as the minimum amount of energy required to remove the outermost electron from an isolated gas atom
- Electron affinity is defined as the amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uni-negative gaseous ion.
- Electronegativity is the ability of an atom to attract the electrons towards itself in a chemical bond

References for Further Information

- B.Earl and LDR Wilford, Introducion to Advanced Chemistry.
- Iain Brand and Richard Grime, Chemistry (11-14).
- Lawarie Ryan, Chemistry for you.

Exercise

1. Choose the correct answer

(i) In the periodic table the number of period of an element is same as

(a) atomic number (b) atomic mass

(c) group number (d) principal quantum number

(ii) Possess highest electronegativity

(a) F (b) Cl

(c) Br (d) I

(iii) Which of the following will produce a base in water?

(a) Na_2O (b) CO_2

(c) SO_3 (d) SiO_2

(iv) Which is correct for $_{12}Mg$?

(a) period 1 element (b) group 1 elemetn

(c) period 2 element (d) alkaine earth metal

(v) An element has electron configuration $1s^2, 2s^2, 2p^6, 3s^2, 3p^2$, it belongs to:

(a) period 2 (b) group 14

(c) group IIIA (d) period 1

(vi) Which of the following element has highest ionization energy?

(a) Li (b) Na

(c) K (d) Rb

(vii) Which of the following element has lowest electron affinity?

(a) F (b) Cl

(c) Br (d) I

(viii) Which of the following oxide is amphoteric?

(a) MgO (b) CO_2

(c) Al_2O_3 (d) SO_3

(ix) Element of which group have strong non-metallic character?

(a) G-IA (b) G-IIIA

(c) G-IVA (d) G-VIIA

(x) Which atom has smaller ionization energy?

(a) B (b) Al

(c) C (d) N

2. Give short answer.

(i) What is a period?

(ii) Differentiate between s and p-block elements.

(iii) Fine out the position of sulphur (atomic number 16) in the periodic table.

(iv) Find out the position of an element in the periodic table, which has electronic configuration: $1s^2,2s^2,2p^6,3s^2,3p^3$

(v) In what respect metals differ from metalloids?

(vi) A cation is smaller than its atom. Justify

3. Describe the varication of electron affinity in the periodic table.

4. Explain the reactions of sodium and magnesium with
   (i) Oxygen
   (ii) Chlorine
   (iii) Water
5. Discuss the variation in the nature of oxides of period 3 elements.
6. Explain the variation in metallic and non-metallic character in the periodic table.
7. Evaluate on the following giving reasons:
   (i) Atomic radii decrease in a period
   (ii) Atomic radii increase in a group
8. How does the periodic table help in understanding the properties of elements?
9. Evaluate the significance of the periodic table.
10. How does periodic table aid in understanding the characteristics of oxides across the period in the periodic table.

Project:

Assign each student an element from the periodic table (Group 1 and Group 2) have them create a card, each card should indicate information about the element's properties, uses, and interesting facts.