

PERIODIC TABLE AND PERIODICITY OF PROPERTIES

Student Learning Outcomes (SLOs)

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

- Define the periodic table as an arrangement element in periods and group, in order of increasing proton number/atomic number.
- Identify the group or period or block of an element using its electronic configuration (only the idea of subshells related to the blocks can be introduced).
- Explain the relationship between group number and the charge of ions formed from elements in the group in terms of their outermost shells.
- Explain similarities in the chemical properties of elements in the same group in terms of their electronic configuration.
- Identify trends in groups and periods, given information about the elements, including trends for atomic radius, electron affinity, electronegativity, ionization energy, metallic character, reactivity, and density.
- Use terms like alkali metals, alkaline earth metals, halogens, noble gases, transition metals, lanthanides, and actinides in reference to the periodic table.

- Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity.
- Deduce the nature, possible position in the Periodic Table and the identity of unknown elements from given information about their physical and chemical properties.
- Define Group 1 Alkali metals as relatively soft metals with general trends down the group limited to decreasing melting point, increasing density and increasing reactivity.
- Predict properties of other elements in Group 1, given information about the elements.
- Predict properties of elements in Group 1 in order of reactivity given relevant information.
- Define Group VII halogens as diatomic non-metals with general trends limited to increasing density and decreasing reactivity.
- Identify the appearance of halogens at rtp as fluorine as pale yellow gas, chlorine as yellow-green gas, bromine as red-brown liquid, iodine as grey-black solid.
- Explain the displacement reactions of halogens with other halide ions and also as reducing agents.
- Predict the properties of elements in group VII, given information about the elements.
- Analyze the relative thermal stabilities of the hydrogen halides and explain these in terms of bond strengths.
- Describe the transition elements as metals that: have high densities, high melting points, variable oxidation numbers, form coloured compounds and act as catalysts for industrial purposes. (some example include catalysts being used are the Haber process, catalytic converters, Contact process and manufacturing of margarine).
- Define the Group 18 noble gases as un-reactive mono-atomic gases.
- Explain this in terms of electronic configuration.
- Compare the general physical properties of metals and non-metals. (specifically in terms of
 - a. Thermal conductivity
 - b. Electrical conductivity
 - c. Malleability and ductility
 - d. Melting points and boiling points

INTRODUCTION

Welcome to the exciting world of chemistry, where the elements come to life thanks to the remarkable periodic table. From its humble beginnings, where only 23 elements were known until the end of the 18th century, to its development of 118 elements today. It is very difficult and impossible to remember information about the reactions, properties, and atomic masses of elements. So we obviously need a way to organize our information about them. The periodic table is one of the most important tools in chemistry. It is very useful for understanding and predicting the properties of elements. For example, if you know the physical and chemical properties of one element in a group, you can predict the physical and chemical properties of any other element in the same group. The periodic table allows you to relate the reactivity tendencies of elements to their atomic structure. You can also predict which elements can form ionic or covalent bonds.

4.1 PERIODIC TABLE

One of the most important activities is the search for order. A large number of observations or objects can be arranged into groups according to common features they share, it becomes easier to describe them. After the discovery of atomic number by Moseley in 1913, it was noticed that atomic number could serve as a base for systematic arrangement of elements. Thus elements are

arranged in the order of increasing atomic number. A table showing systematic arrangement of elements is called **periodic table**. It is based on the Periodic law that states if the elements are arranged in the order of their increasing atomic numbers, their properties are repeated in a periodic manner.

4.1.1 Periods and Groups of Elements.

The most commonly used form of the periodic table is shown in figure 4.1. Note that the elements are listed in order of increasing atomic numbers, from left to right and from top to bottom. Hydrogen (H) is in the top left corner. Helium (He), atomic number 2, is at the top right corner. Lithium (Li), atomic number 3, is at the left end of the second row.

The horizontal rows of the periodic table are called **periods**. There are varying number of elements in periods. How many periods you find in the periodic table? There are seven periods. The number of elements per period range from 2 in period 1 to 32 in period 6. First three periods are called short periods and the remaining periods are called long periods. The properties of elements within a period change gradually as you move from left to right in it. But when you move from one period to the next, the pattern of properties within a period repeats. This is in accordance to the periodic law.

International Union of Pure and Applied Chemistry (IUPAC) has recently renamed newly discovered elements and placed them on the periodic table.

Activity 4.1

Look at the periodic table and write number of elements present in the relevant period in the table

Table Number of elements in the periods of the periodic table	
Period No.	No. of elements
First	
Second	
Third	
Fourth	
Fifth	
Sixth	
Seventh	

Figure 4.1: Periodic Table of Elements

Atom number — 1 — Symbol — H — Name — Hydrogen — Atomic weight — 1.008 — Electrons per shell — 1

Subcategory metals, nonmetals, and metalloids

- Alkali metals
- Alkaline earth metals
- Transition metals
- Lanthanides
- Actinides
- Post transition metals
- Metalloids
- Reactive non metals
- Noble gases
- Unknown properties

1 IA H Hydrogen 1.008 1	2 IIA He Helium 4.003 2	3 Li Lithium 6.941 3	4 Be Beryllium 9.012 4	5 B Boron 10.81 5	6 C Carbon 12.01 6	7 N Nitrogen 14.01 7	8 O Oxygen 16.00 8	9 F Fluorine 18.99 9	10 Ne Neon 20.18 10	11 Na Sodium 22.99 11	12 Mg Magnesium 24.31 12	13 Al Aluminum 26.98 13	14 Si Silicon 28.09 14	15 P Phosphorus 30.97 15	16 S Sulfur 32.06 16	17 Cl Chlorine 35.45 17	18 Ar Argon 39.95 18	19 K Potassium 39.10 19	20 Ca Calcium 40.08 20	21 Sc Scandium 44.96 21	22 Ti Titanium 47.88 22	23 V Vanadium 50.94 23	24 Cr Chromium 52.00 24	25 Mn Manganese 54.94 25	26 Fe Iron 55.85 26	27 Co Cobalt 58.93 27	28 Ni Nickel 58.69 28	29 Cu Copper 63.55 29	30 Zn Zinc 65.38 30	31 Ga Gallium 69.72 31	32 Ge Germanium 72.64 32	33 As Arsenic 74.92 33	34 Se Selenium 78.96 34	35 Br Bromine 79.90 35	36 Kr Krypton 83.80 36	37 Rb Rubidium 85.47 37	38 Sr Strontium 87.62 38	39 Y Yttrium 88.91 39	40 Zr Zirconium 91.22 40	41 Nb Niobium 92.91 41	42 Mo Molybdenum 95.94 42	43 Tc Technetium [98] 43	44 Ru Ruthenium 101.07 44	45 Rh Rhodium 102.91 45	46 Pd Palladium 106.42 46	47 Ag Silver 107.87 47	48 Cd Cadmium 112.41 48	49 In Indium 114.82 49	50 Sn Tin 118.71 50	51 Sb Antimony 121.76 51	52 Te Tellurium 127.60 52	53 I Iodine 126.91 53	54 Xe Xenon 131.29 54	55 Cs Cesium 132.91 55	56 Ba Barium 137.33 56	57-71 Ln Lanthanides [57-71] 57-71	72 Hf Hafnium 178.49 72	73 Ta Tantalum 180.95 73	74 W Tungsten 183.84 74	75 Re Rhenium 186.21 75	76 Os Osmium 190.23 76	77 Ir Iridium 192.22 77	78 Pt Platinum 195.08 78	79 Au Gold 196.97 79	80 Hg Mercury 200.59 80	81 Tl Thallium 204.38 81	82 Pb Lead 207.2 82	83 Bi Bismuth 208.98 83	84 Po Polonium [209] 84	85 At Astatine [210] 85	86 Rn Radon [222] 86	87 Fr Francium [223] 87	88 Ra Radium [226] 88	89-103 Ac Actinides [89-103] 89-103	104 Rf Rutherfordium [261] 104	105 Db Dubnium [262] 105	106 Sg Seaborgium [266] 106	107 Bh Bohrium [264] 107	108 Hs Hassium [277] 108	109 Mt Meitnerium [268] 109	110 Ds Darmstadtium [271] 110	111 Rg Roentgenium [272] 111	112 Cn Copernicium [285] 112	113 Nh Nihonium [284] 113	114 Fl Flerovium [289] 114	115 Mc Moscovium [288] 115	116 Lv Livermorium [293] 116	117 Ts Tennessine [294] 117	118 Og Oganesson [294] 118
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119 Lu Lutetium 174.97 119	120 Yb Ytterbium 173.05 120	121 Tm Thulium 168.93 121	122 Er Erbium 167.26 122	123 Ho Holmium 164.93 123	124 Dy Dysprosium 162.50 124	125 Tb Terbium 158.93 125	126 Gd Gadolinium 157.25 126	127 Eu Europium 151.96 127	128 Sm Samarium 150.36 128	129 Pm Promethium [145] 129	130 Nd Neodymium 144.24 130	131 Pr Praseodymium 140.91 131	132 Ce Cerium 140.12 132	133 La Lanthanum 138.91 133	134 Ac Actinium [227] 134	135 Th Thorium 232.04 135	136 Pa Protactinium 231.04 136	137 U Uranium 238.03 137	138 Np Neptunium 237.05 138	139 Pu Plutonium 244.06 139	140 Am Americium 243.06 140	141 Cm Curium 247.07 141	142 Bk Berkelium 247.07 142	143 Cf Californium 251.08 143	144 Es Einsteinium 252.08 144	145 Fm Fermium 257.10 145	146 Md Mendelevium 258.10 146	147 No Nobelium 259.10 147	148 Lw Lawrencium [260] 148
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Elements that have similar properties lie in the same column in the periodic table. Each vertical column of elements in the periodic table is called a **group** or **family**.

Two numbering systems are often used to designate groups. You should know both. 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) decided that the groups would be 1-18 from left to right.

The elements in the same group have same number of valence electrons. Group number indicates the number of valence electrons in an element. For example, Group 1 and Group 2 elements have 1 and 2 valence electrons respectively. In Groups 13 elements have 3, Group 14 have 4, Group 15 have 5 valence electrons and so on. It is important to note that in Groups 13 to 18 (p block elements), the number of valence electrons is equal to group number minus 10.

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

Names of Some Groups in the Periodic Table

Some groups of elements in the periodic table have been given group names. For example metallic elements in Group 1 are called alkali metals. Group 2 elements are called alkaline earth metals. The elements in Group 17 or VIIA are called halogens. The Group 18 or VIIIA elements are called noble gases because they do not readily undergo chemical reactions.

Recall that all elements have a unique identification number known as the atomic number or proton number. The atomic number of an element represents the number of electrons or protons present in the atom of the element. Aufbau's Principle helps in determining the order in which the electron orbitals get filled.

Electronic Configuration

According to Aufbau's principle, the order in which the orbitals fill up is as follows:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p and so on.

Each orbital has a fixed capacity for the maximum number of electrons accommodated, s-orbitals have the capacity of 2 electrons, while p orbitals have the capacity for 6 electrons, d orbitals have the capacity for 10 electrons and f orbitals have the capacity for 14 electrons.

Using these concepts, we can determine the electronic configuration of the given element.

Block of an element: When you have filled all the electrons, the orbital in which the last electron is in, represents the block in which the element is placed.

Period of an element: Now, to determine the period in which the element is placed, you need to look at the principal quantum number (n) of the valence electron. This number represents period number of element

Group of an element: To determine the group, we need to understand some rules:

- If the element is in s block, then the group number is equal to the number of valence electrons.
- If the element is in the p block, then the number of the group can be determined by the formula: (number of valence electrons + 10).

For example, the atomic number of sodium is 11.

Hence its electronic configuration is: $1s^2, 2s^2, 2p^6, 3s^1$

Since the valence electron is in the 3s subshell, sodium belongs to the s block.

The principal quantum number of the valence electron of Na is 3. Hence, it belongs to the 3rd period.

Since Na belongs to the s block, its group number is equal to a number of electrons in valence subshell s. This is equal to 1

Hence, sodium belongs to the Group 1.

Note:

we can start filling the orbitals in the order mentioned by the Aufbau principle.

$1s, 2s, 2p, 3s, 3p, 4s$

Example 4.1: Identifying the group and period of an element

Identify the group, period, and block of following elements on the basis of electronic configuration.

1. Al (atomic number = 13)
2. K (atomic number = 19)

Problem Solving Strategy:

Write electronic configuration of element. Identify its valence shell. Remember that n value of the valence shell indicates period. Total number of electrons in the valence shells represents group number if element belongs to s block. If it belongs to p block, then group number is equal to the total number of valence + 10.

Solution:

1. Electronic configuration of Al (atomic no. 13) = $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

Valence sub-shells is 3p, so Al belongs to p block

As $n = 3$, Al is present in the 3rd period.

Total number of electrons in the valence shell = $2 + 1 = 3$

Group number of Al = total number of electrons in the valence sub-shells + 10

$$= 3 + 10$$

$$= 13$$

Hence Al belongs to Group 13

2. Electronic configuration of K (atomic no. 19) = $1s^2, 2s^2, 2p^6, 3s^2, 3p^4, 4s^1$

Valence shells is 4s. hence K belongs to s block

As $n = 4$, K is present in the 4th period.

Total number of electrons in the valence shell = 1

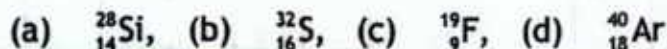
Group number of K = total number of electrons in the valence sub-shells

$$= 1$$

Hence K belongs to Group 1

CONCEPT ASSESSMENT EXERCISE 4.2

Identify the group and period of the following elements on the basis of electronic configurations.



Example 4.2: Classifying or dividing elements into groups and periods

Electronic configuration of atoms of some elements are given below. Classify them in groups and periods.

- A. $1s^2 2s^2$
- B. $1s^2 2s^2 2p^3$
- C. $1s^2 2s^2 2p^5$
- D. $1s^2 2s^2 2p^6 3s^2$
- E. $1s^2 2s^2 2p^6 3s^2 3p^5$
- F. $1s^2 2s^2 2p^6 3s^2 3p^3$

Problem solving Strategy:

Remember that:

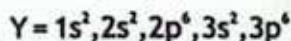
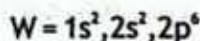
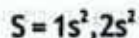
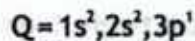
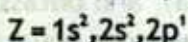
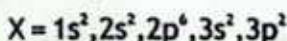
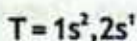
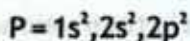
- The elements whose atoms have similar valence shell electronic configuration belong to the same group.
- The n value of the valence shell indicates period.
- The elements whose atoms have same value of n for the valence shell lie in the same period.

Solution:

	II A	V A	VII A	
2	A $2s^2$	B $2s^2 2p^3$	C $2s^2 2p^5$	Period
3	D $3s^2$	F $3s^2 3p^3$	E $3s^2 3p^5$	Period

CONCEPT ASSESSMENT EXERCISE 3.1

Electronic configuration of atoms of some elements are given below. Place them into groups and periods.



IA		VIIIA						
	IIA							
			IIIA	IVA	VA	VIA	VIIA	

4.1.2 s and p Blocks in the Periodic Table

Group 1 and Group 2 elements contain their valence electrons in the s sub-shell. Therefore, these elements are called s-block elements. Elements in groups 13 to 18 (except He) are known as p-block elements because their valence electrons are located in the p sub-shell. Lanthanides and actinides are known as f-block elements since their valence electrons lie in f sub-shell. Figure 4.2 shows the blocks of the periodic table.

$\text{Li} = 1s^2, 2s^1$, as valence electron is in s sub-shell, Li belongs to s-block.

$\text{C} = 1s^2, 2s^2, 2p^2$, as valence electron is in sub-shell p, C belongs to p-block.

4.2 GROUP NUMBER AND CHARGE ON AN ION

The group number of an element in the periodic table can provide information about the charge of an ion formed by an element. Valence electrons are involved in the formation of ions. The relationship between group number and ions formed by elements is based on the number of valence electrons in the element.

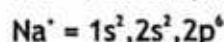
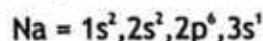
The group number of an s-block element in the periodic table corresponds to its number of valence electrons.

Whereas in the case of p-block elements, the number of valence electrons is equal to Group number minus 10.

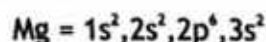
Some elements tend to lose electrons. Why? Elements tend to achieve a stable electron configuration such as the noble gases. Remember that the 2 or 8 electron configuration is the most stable configuration. Elements with 1-3 electrons in their valence shell tend to lose those electrons and form +1, +2, +3 ions respectively. Elements with 5-7 electrons in their valence shell tend to gain 3, 2, 1 electrons respectively and form negatively charged ions with -3, -2, -1 charges respectively. Elements with 4 valence electrons can lose 4 electrons to form +4 ions. They can also gain 4 electrons and form -4 ions.

Group 1 (alkali metals): Group 1 elements such as lithium (Li), sodium (Na), and potassium (K) have one valence electron and belong to s block. S block elements lose electrons equal to their group number. They tend to lose this electron to form a +1 ion, also known as a mono-valent cation. For example: Lithium (Li) loses one valence electron to form Li^+ . Sodium (Na) loses one

valence electron to form Na^+ . Potassium (K) loses one valence electron to form K^+ . These elements after losing an electron acquire 8 electron configuration of a noble gas.

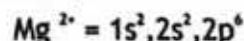


Group 2 (alkaline earth metals): Group 2 elements such as beryllium (Be), magnesium (Mg), and calcium (Ca) have two valence electrons and are s block element. They tend to lose these two electrons to form + 2 ions, also called divalent cations. For example: Beryllium (Be) loses two valence electrons to form Be^{2+} . Magnesium (Mg) loses two valence electrons to form Mg^{2+} . Calcium (Ca) loses two valence electrons to form Ca^{2+} .

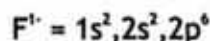
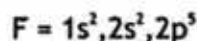


How many electrons Mg can lose to achieve stable electron configuration?

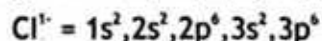
Magnesium will lose 2 electrons to achieve stable configuration and this no. is same as its group number i.e., 2



Some elements tend to gain electrons to achieve noble gas configuration. For example, Group 17 (Halogens): Group 17 elements such as fluorine (F), chlorine (Cl), and bromine (Br) have seven valence electrons. They tend to gain one electron to reach a stable octet and form - 1 ion, also called a monovalent anion. For example: Fluorine (F) gains one electron to form F^- . Chlorine (Cl) gains one electron to form Cl^- . Bromine (Br) gains one electron to form Br^- .



Similarly, $\text{Cl} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$



Group 16 (chalcogens): Group 16 elements such as oxygen (O), sulfur (S), and selenium (Se) have six valence electrons. They tend to gain two electrons to reach a stable octet and form a -2 ion, also called a divalent anion. For example: Oxygen (O) gains two electrons to form O^{2-} . Sulfur (S) gains two electrons to form S^{2-} .

Group 18 (precious gases): Group 18 elements such as helium (He), neon (Ne), and argon (Ar) have full valence electron shells (except helium, which has only two valence electrons). They are chemically stable and do not form ions under normal conditions. Noble gases are known for their low reactivity due to their stable electronic configuration.

Example 4.3: Obtaining the position of element in the periodic table from the electronic configuration

Find out the position of the following elements in the periodic table from the electronic configuration:

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

Problem Solving Strategy:

Write electronic configuration of the element. Identify the valence shell configuration, coefficient of s or p sub-shell represents period number and total number of electrons in 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
 Number of valence electrons = $2 + 3 = 5$
 N belongs to p-block
 So, Group number = $5 + 10 = 15$
 Nitrogen is present in the 2nd period of 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
 Number of valence electrons = $2 + 4 = 6$
 O belongs to p-block
 So, Group number = $6 + 10 = 16$
 Oxygen is present in the 2nd period of Group 16

Figure 4.2: Periodic Table of Elements

s-block

ns^{1 to 2}

24

Cr

Chromium

[Ar] 3d⁵ 4s¹

Atomic number

Symbol

Electronic configuration

p-Block

ns² np^{1 to 6}

d-Block

(n-1)d^{1 to 10} ns^{0, 1 or 2}

f-Block

(n-2)f^{1 to 14} (n-1)d^{0 or 1} ns²

1 H																	2 He																								
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne																								
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar																								
19 K	20 Ca											21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr														
37 Rb	38 Sr											39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe														
55 Cs	56 Ba											57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra											89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Obtain the valence shell configuration of Al and S from their position in the periodic table

4.3 PERIODICITY OF PROPERTIES

There is a periodic fluctuation in the electronic configuration of the elements as the atomic number increases. Therefore, the physical and chemical properties of the elements vary in a periodic manner. Elements with a similar valence shell electronic configuration are placed in the same group, one below the other. Chemical properties depend on the electronic configuration of the valence shell. Because all the elements in a given group have similar valence shell electronic configurations, 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. 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. In this section, you will learn about the variation of physical properties of certain elements within a group and across a period.

4.3.1 Shielding Effect

Figure 4.3 shows electronic configuration of Li, Be and Mg.

Which atom has more shells, Be or Mg? Which atom has more electrons between the nucleus and the valence electrons, Be or Mg?

Electrons present in the inner shells cut off attractive force between the nucleus and the valence electrons.

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.

Which atom has greater shielding effect, Be or Mg?

As you move from top to bottom in a group the number of electronic shells increase. So the number of electrons in the inner shell also increase. As a result shielding effect increases.

Which atom, Li or Be has greater number of shells? Which atom, Li or Be has greater number of electrons between nucleus and valence electrons?

As you move from left to right in a period the number of electrons in the inner shells remains constant, therefore, shielding effect remains constant.

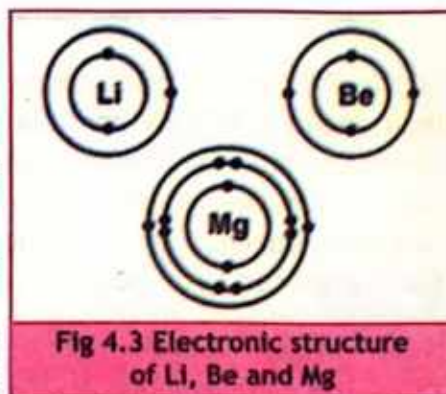


Fig 4.3 Electronic structure of Li, Be and Mg

Example 4.4: Identifying the element whose atoms have greater shielding effect, using periodic table

Choose the elements whose atoms you expect to have greater shielding effect.

- (a) Be or Mg (b) C or Si

Problem Solving Strategy:

Look at the periodic table and find the relative position of given elements in the periodic table. Apply the trend of increasing shielding effect in a group.

Solution:

- (a) Mg atoms will have greater shielding effect.
(b) Si atoms will have greater shielding effect.

CONCEPT ASSESSMENT EXERCISE 4.5

Choose the element whose atoms you expect to have smaller shielding effect.

- (a) F or Cl (b) Li or Na (c) B or Al

All the physical and chemical properties of elements depend on the electronic configuration of their atoms. We now consider some properties of atoms that are affected by electronic configuration: atomic size, ionization energy, electron affinity and electronegativity. They usually increase and decrease repeatedly throughout the periodic table. That is, they show consistent changes or trends within a group or a period. These tendencies are correlated with their behaviour.

4.3.2 Atomic Size

The size of an atom depends on its electronic configuration. **Atomic size is the average distance between the atomic nucleus and the electronic outer shell.** Figure 4.4 shows the atomic radii of the main group elements. Figure 4.4 shows the variation of atomic radii within a period and within a group. You can see two general trends in atomic radii.

- 1) The atomic radius decreases in each period as you move across the period. This is because as you move from one element in the sequence to the next, to the right of it. Another electron is added to the same valence shell. At the same time, the positive charge of the core also increases by one. The attraction of the nucleus to the electron in the valence shell increases. Therefore, the size of the shell

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca	Ga	Ge	As	Se	Br	Kr
Rb	Sr	In	Sn	Sb	Te	I	Xe
Cs	Ba	Tl	Pb	Bi	Po	At	Rn

Figure 4.4: Atomic sizes of the main group elements

and the radius of the atom decreases. For example, going from lithium to beryllium, the atomic size decreases. You can understand this from the electronic configuration of the valence shell of Li ($2s^1$) and B ($2s^2$). Moving from Li to Be, the number of shells does not change, but the atomic number increases from 3 to 4. Therefore, the strength of the nucleus on the valence shell electron increases. Therefore, the atomic radius decreases.

- 2) Atomic radius increases in each main group as you move down the element group. This is because the size of an atom is determined by the size of its valence shell. As you move down the group to the next lower element, the atom has an additional shell of electrons. This increases the radius of the atom. For example, going from Li to Na, the atomic radius increases. Consider the electron configuration of Li ($1s^2 2s^1$) and Na ($1s^2, 2s^2, 2p^6, 3s^1$). A new electron shell is added, increasing the size of the atom.

Example 4.5: Identifying the element that has greater atomic radius

Choose the element whose atom you expect to have larger atomic radius in each of the following pairs.

- (a) Mg, Al (b) C, Si

Problem Solving Strategy:

Remember that the larger atom in any:

- (a) Period lies further to the left in the periodic table.
- (b) Group lies closer to the bottom in the periodic table.
- (c) Check the periodic table and choose the element.

Solution:

- (a) The larger atom is Mg
- (b) The larger atom is Si

CONCEPT ASSESSMENT EXERCISE 4.6

Using the periodic table but without looking at the figure 4.4, choose the element whose atom you expect to have smaller atomic radius in each of the following pairs.

- (a) O or S (b) O or F

4.3.3 Ionization Energy

Ionization energy is an important property of atoms that explains cation formation. "Ionization energy is defined as the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom".



	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	H 1312							He 2372
2	Li 520	Be 899	B 801	C 1086	N 1402	O 1314	F 1681	Ne 2081
3	Na 496	Mg 738	Al 578	Si 786	P 1012	S 1000	Cl 1251	Ar 1521
4	K 419	Ca 590	Ga 579	Ge 762	As 947	Se 941	Br 1140	Kr 1351
5	Rb 403	Sr 549	In 558	Sn 709	Sb 834	Te 869	I 1008	Xe 1170
6	Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At 926	Rn 1037

Figure 4.5 Ionization energies of the main group elements

Ionization energy is a measure of the extent to which the nucleus attracts the outermost electron. A high value of ionization energy means stronger attraction between the nucleus and the outermost electron. Whereas a low ionization energy indicates a weaker force of attraction between the nucleus and the outermost electron. Figure 4.5 shows the ionization energies of the main group elements. Values are given in units of kJ mole^{-1} or kJ/mole .

Trends in ionization energy values.

The value of the ionization energy decreases from top to bottom in the group. This is because the shielding effect of the atoms increases down the group. Greater shielding effects result in a weaker attraction of the valence electrons to the nucleus. So they are easier to remove. This leads to a decrease in ionization energy from top to bottom in the group. Which atom has a greater shielding effect, Li or Na? As you move from left to right in the period, the shielding effect remains unchanged. But little by little the nuclear charge increases. The stronger attraction between the nucleus and the valence electron increases. As a result, the ionization energy increases from left to right in a period. Which atom has the higher ionization energy, Li or Be?

Example 4.6: Identifying the element that has smaller ionization energy

Choose the element whose atom you expect to have smaller ionization energy in each of the following pairs.

- (a) B, C (b) N, P

Problem Solving Strategy

Remember that ionization energy:

- (a) Increases across a period. The element that has smaller ionization energy will be further to the left in the periodic table.

- (b) Decreases from top to bottom in a group. The element that has smaller ionization energy will correspond to the element closer to the bottom.
- (c) Check the periodic table to choose the element.

Solution:

- (a) The atom with the smaller ionization energy is B.
- (b) The atom with the smaller ionization energy is P.

CONCEPT ASSESSMENT EXERCISE 4.7

Which atom has the smaller ionization energy?

- (a) B or N (b) Be or Mg (c) C or Si

4.3.4 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.

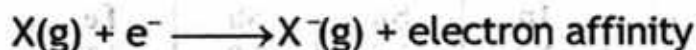


Figure 4.6 shows electron affinities of main group elements.

Factors affecting electron affinity are nuclear charge, atomic radius and shielding effect.

As you move from left to right through a period, electron affinity generally increases. This is due to an increase in nuclear charge and a decrease in atomic radius, which binds the extra electron more tightly to the nucleus. But the shielding effect remains constant in each cycle. Therefore, the alkali metals have the lowest and the halogens the highest electron affinities in each period.

Electron affinity decreases from top to bottom in a group. This is due to an increase in the shielding effect. Due to the increased shielding effect and increase in atomic radius, the added electron binds less tightly to the nucleus. As a result, less energy is released.

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 4.6 electron affinities of main group elements

There are several exceptions to the general trend of election affinity values. You will learn reasons for it in higher grade.

4.3.5 Electronegativity

Electronegativity is the ability of an atom to attract electrons toward itself in a chemical bond. Figure 4.7 shows as a scale of electronegativities of the elements devised by Linus Pauling. The American chemist Linus Pauling devised a method for calculating the relative electronegativities of elements.

H 2.1								He
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0		Ne 2.1
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0		Ar 3.0
K 0.5	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8		Kr 2.1
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5		Xe 2.6
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2		Rn
Fr 0.7	Ra 0.9							

Figure 4.7 the electronegativities of elements.

Activity 4.2

Determining the general trends in the electronegativities

You will need:

- Figure 4.7

Carry out the following:

1. Move across the second period from left to right and note down the variation in electronegativity values.
2. Move across the 3rd period from left to right and note down the variation in electronegativity values.
3. Make generalization about the variation in electronegativities across a period and write reason.
4. Move from top to bottom in Groups IA and IIA and note down the variation in electronegativities value.

5. Move from top to bottom in Groups VIA and VIIA and note down the variation in electronegativities value.

Make generalization about the trend in electronegativity values in a group. Give reason.

4.4 CHARACTERISTIC PROPERTIES

Characteristic properties of an element in a given group are based on periodicity and chemical reactivity. For example, in Group 1 (alkali metals) such as lithium, sodium, potassium, are highly reactive metals. They have general electron configuration ns^1 . Their reactivity trend increases as you move down the group. Lithium, being at the top of the group is the least reactive metal among alkali metals. As you move down the group, the atomic size increases and the outer most electron is further from the nucleus, leading to be lost easily. This leads to increased reactivity. So, sodium is more reactive than lithium. Which is more reactive sodium or potassium?

Similar trend is observed in Group 2 (alkaline earth metals). Which is more reactive Mg or Ca?

4.4.1 Metallic Character

Metallic nature refers to a property of elements in the periodic table that determines how easily they can lose electrons and form positive ions (cations). Elements with metallic character have a strong tendency to lose electrons and easily form cations. The metallic character of an element is affected by its position in the periodic table.

Metallic character increases as you move down a group in the periodic table. This is primarily due to addition of new electronic shells. The outermost electrons are farther from the nucleus and experiences weaker attractive forces, making it easier for them to be lost. This promotes metallic character.

Metallic character decreases as you move across a period from left to right in the periodic table. This is because effective nuclear charge increases across a period, while the number of shells remains the same. The stronger attractive forces make it more difficult for valence electrons to be lost.

Example 4.6: Identifying the element that has higher metallic character.

Choose the element you expect to have higher metallic character in each of the following pairs.

- (a) Na or K
- (b) Na or Mg

Problem Solving Strategy

Remember that metallic character :

- (a) Increases down the group. The element that has higher metallic character will be closer to the bottom.
- (b) Decreases across a period. The element that has higher metallic character is further to the left.
- (c) Check the periodic table to choose the element.

Solution:

- (a) K
- (b) Mg

CONCEPT ASSESSMENT EXERCISE 4.8

Which element has lower metallic character?

- (a) Li or K
- (b) Mg or Ca
- (c) Compare and contrast ionization energy and electron affinity

4.4.2 Reactivity

The capability of an element to react with other elements to form new compounds is called it reactivity. Reactivity of elements generally increases as you move down a group. This is due to the increase in atomic size. The outermost electrons are farther from the nucleus and experience weaker attractive forces, making it easier for them to participate in chemical reaction.

Reactivity tends to vary across a period. Elements on the left side of a period (Group 1 and 2) are highly reactive due to their strong tendency to lose electrons and form positive ions. Elements on the right side of a period (Group 16 and 17) are highly reactive as well but tend to gain electrons to form negative ions.

4.4.3 Density

Density of elements generally increases as you move down a group. This is due to the increasing atomic mass and the larger size of atoms. As the number of protons and neutrons in the nucleus increases, the atomic mass increases. This results in higher density.

Density can vary across the period. In general, density tends to increase from left to right until it reaches a maximum around the middle of the period, and then it starts to decrease.

4.4.4 Characteristic Properties of Alkali Metals

Some characteristic properties of alkali metals are as follows:

- (a) **Highly reactive metals:**
Alkali metals are highly reactive metals in the periodic table. They readily lose valence electron to form a +1 cations. This trend increases down the group.
- (b) **Softness and low density:**
Alkali metals have low densities and are relatively soft, which allows them to be easily cut with a knife. This trend increases down the group the group.
- (c) They are excellent conductors of electricity and heat.
- (d) They have low melting points.
- (e) They are highly reactive and monovalent elements.
- (f) They react with H_2O to give H_2 and alkali metal hydroxides.

Which is more soft Na or K?

4.4.5 Prediction of properties of other elements in Group 1

In Group 1 lithium, sodium and potassium are a collection of relatively soft metals showing a trend in melting point and reaction with water.

The metals in group I are called alkali metals.

- They are very soft.
- Their melting and boiling points decrease down the group.
- When alkali metals react with water, they produce a metal hydroxide and hydrogen.

$$\text{metal} + \text{water} \longrightarrow \text{metal hydroxide} + \text{hydrogen}$$
- The alkali metals become more reactive down the group.

Activity 4.5

Predict the properties of other elements in Group I, from the data given above.

The element after Potassium is Rubidium and you can predict that its reaction with water will be much more violent. We can also predict that Rubidium will have a lower melting and boiling point than the three elements above it. And the elements below Rubidium will be even more reactive and have very low melting and boiling points. It will also react with water to form metal hydroxide and hydrogen.

4.4.6 Position of Unknown Element in the Periodic Table

You can place an unknown element accurately at appropriate position in the periodic table, and can predict about its properties.

The electronic configuration of an element strongly influences its chemical behaviour. Elements within the group have similar electronic configuration, and therefore similar properties. By examining the electronic configuration of unknown element and comparing it to the known elements in the periodic table, its likely position in the periodic table can be determined.

The periodic table allows for the identification of trends and patterns across periods and groups. These trends include variations in atomic size, ionization energy, electron affinity, electronegativity, reactivity, and other properties. By analyzing these trends, it becomes possible to estimate the general properties of unknown element and make guess about its position in the periodic table.

Example 4.7: Identifying position of an unknown element in the periodic table

Suppose you have an unknown element having atomic number 19, 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 : $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$

Valence shell electronic configuration is $4s^1$, which shows it is an alkali metal, because, Group 1 elements have one electron in valence sub-shell s . As n value of valence sub-shell is 4, this unknown element must lie in the 4th period in the periodic table.

From its position in the periodic table, you can predict its properties. For example will it possess higher or lower melting point, density, reactivity, etc. than the element above or below it.

4.5 TRANSITION ELEMENTS

Elements located in d-block (Group 3 to 12) in the periodic table are called transition elements. These elements exhibit several characteristic properties, which set them apart from other elements. Some of their properties are as follows.

1. High Density

Transition elements generally possess high densities due to their higher atomic masses and closely packed structures. For example, iron(Fe) has a density of 7.87 g/cm^3 , tungsten(W) has a density of 19.3 g/cm^3 .

2. High Melting Points

Transition elements have high melting points. This is because their metallic bonding is stronger, which in turn is due to the presence of partially filled d-sub shells. For example tungsten has a melting point of 3422°C , platinum(Pt) has a melting point of 1768°C .

3. Variable Oxidation States

Transition elements exhibit multiple oxidation states. This is because of d-sub shell can also participate in bonding along with s-sub shell. For example, iron(Fe) has oxidation states +2 and +3, copper(Cu) has oxidation states +1 and +2.

4. Coloured Compounds

Transition elements often exhibit vibrant colours. For example, copper compounds appear blue or green, chromium compounds are often red or green.

5. Catalysts for Industrial Processes

Transition metals and their compounds are widely used as catalyst in various industrial processes. For example,

- Iron is used in the Haber Process for the synthesis of ammonia.
- Platinum and palladium are used as catalyst in catalytic converters to reduce harmful emissions in automobiles and industrial units.
- Nickel is used as catalyst in the manufacture of margarine.
- Platinum is used as catalyst in the contact process for the manufacture of sulphuric acid.

4.6 LANTHANIDES & ACTINIDES

Lanthanides also known as "rare earth elements" are series of 14 elements located at the bottom of the periodic table. They include elements with atomic number 57 to 71.

Actinides are another series of 14 elements located just below lanthanides. They include elements with atomic number 89 to 103.

4.7 HALOGENS

The elements in group 17 (or Group VII-A) are called halogens. The name halogen is derived from the Greek words “halous” meaning salt and “gen” meaning former. Halogens include fluorine, chlorine, bromine, iodine, astatine, and tenessine (astatine and tenessine are radio- active elements. Little is known about their properties). All halogens are reactive non-metals and exist as diatomic molecules.

4.7.1 Appearance of halogens

They all exist as diatomic coloured molecular substances. The colour of halogen become darker as you go down the group. At room temperature and pressure(RTP) fluorine(F_2) exist as pale yellow gas, chlorine(Cl_2) as yellow-green gas, bromine(Br_2) as red-brown liquid and iodine(I_2) as grey-black solid. Iodine easily turn into a dark purple vapours on warming.

Electronic Configuration

Halogens possess 7 electrons in their valence shell. They have general electronic configuration ns^2np^5 . They need only one electron to complete their valence shell. Consequently, they tend to gain one electron to form univalent negative ions, F^- , Cl^- , Br^- , I^- .

Density of halogens

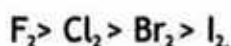
As you move down the group the number of electrons and protons increases, the size of the atom increases and the volume increases. However, the increase in mass exceeds the increase in volume, so the density, which is mass per unit volume, increases in general. Also fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid . So, the forces of attraction between molecules increase down the group. Solid iodine has molecules that are highly attracted and tightly packed together than bromine. Therefore, as you go down the group of halogens, the forces of attraction between molecules increase and the density of the halogen increases.

Densities of halogens

Halogen	Density (g/cm ³ at 25 °C)
Fluorine	0.0017
Chlorine	0.0032
Bromine	3.1028
Iodine	4.933

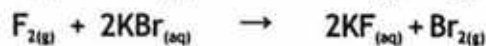
Reactivity of halogens

The reactivity of halogens is directly related to their ability to gain an electron and form a halide ion (fluoride ion F^- , chloride ion Cl^- , bromide ion Br^- , iodide ion I^-) when they react with other elements. Fluorine has the greatest tendency to gain electrons and form a halide ion, making it the most reactive halogen. As you move down the group, the electronegativity of the halogens decreases. This leads to a decrease in reactivity. Which halogen is the least reactive? Bromine or iodine. Because halogens have a strong tendency to gain electrons, they have a strong oxidizing power, and this power decreases down the group. Thus, the order of decreasing oxidizing power is.



Displacement reactions of halogens

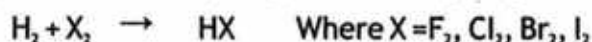
Oxidizing power of F_2 is the highest and that of I_2 is lowest. Due to the relative strength as oxidizing agent, it is possible for a free halogen to oxidize or displace the ion of halogen next to it in the group from their aqueous solutions. This means F_2 can oxidize and displace all the halide ions to free halogen. For example,



Similarly Cl_2 can oxidize Br^- and I^- ions. But I_2 can not oxidize any halide ion.

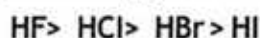
Hydrogen halides and their thermal stabilities

Halogens react with hydrogen to form hydrogen halides.



The strength of the hydrogen-halogen bond is related to the electronegativity difference between the hydrogen and halogen atoms. A larger electronegativity difference results in a stronger bond. As we move from HF to HI, the electronegativity difference between the hydrogen and halogen atoms decreases, resulting in weaker bonds in HCl, HBr, and HI. So, the relative thermal stability of hydrogen halides gradually decreases from HF to HI.

Consequently, the energy needed to break H-X decreases in the following orders



Prediction of properties of elements in Group VIIA or Group 17

The elements present in Group 17 or VIIA are called halogens. They are poisonous non-metals that have low melting and boiling points that increase down the group. As a result of this increasing boiling and melting points, the state of the halogens at room temperature, changes from gas to liquid to solid down the group (fluorine and chlorine, the 1st and 2nd halogens, are a gas; bromine, the 3rd halogen is a liquid; and iodine, the 4th halogen, is a solid). The colours of halogens also get darker from top to bottom.

Activity 4.6

Predict the properties of other elements in Group VII, from the given data given above.

From this data you can predict how the halogens will behave up and down the group. Astatine, the fifth halogen, will have high melting and boiling points so will be solid at room temperature, and will have a very dark colour.

CONCEPT ASSESSMENT EXERCISE 4.9

Which of the following displacement reactions will occur?

- $Cl_{2(g)} + 2NaF_{(aq)} \rightarrow 2NaCl_{(aq)} + F_{2(g)}$
- $Br_{2(l)} + 2KI_{(aq)} \rightarrow 2KBr_{(aq)} + I_{2(s)}$
- $I_{2(s)} + 2KBr_{(aq)} \rightarrow 2KI_{(aq)} + Br_{2(l)}$
- $Cl_{2(g)} + 2KBr_{(aq)} \rightarrow 2KCl_{(aq)} + Br_{2(l)}$
- $Cl_{2(g)} + 2NaI_{(aq)} \rightarrow 2NaCl_{(aq)} + I_{2(s)}$

4.8 NOBLE GASES

Noble gases, also known as inert gases, are a group of chemical elements found in Group 18 (or Group VIII-A) of the periodic table. They have general electron configuration ns^2, np^6 except He, which has $1s^2$. They are characterized by unique properties. They are odorless, colorless mono-atomic gases and possess very low reactivity with other elements. This low reactivity is due to the presence of a complete valence shell, which makes them stable and unlikely to form chemical bonds with other elements under normal conditions. Noble gases include elements: Helium (He) neon (Ne) argon (Ar) krypton (Kr) xenon (Xe) radon (Rn), and oganesson (Og).

DO you know?

Due to their non-reactive nature, noble gases are used in many ways, such as in lighting (e.g. neon signs), refrigeration systems and welding. They are also used in special applications, including filling gas exhaust lines and as a shielding gas in certain industrial processes.

Table 4.1: Electronic Configuration of Noble Gases

Element	Atomic Number	Electronic Configuration
Helium	2	$1s^2$
Neon	10	$1s^2 2s^2 2p^6$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$
Krypton	36	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$
Xenon	54	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6$

4.9 COMPARISON OF GENERAL PHYSICAL PROPERTIES OF METALS AND NON-METALS:

Thermal Conductivity:

Metals generally have high thermal conductivity, which means they can conduct heat easily. On the other hand, non-metals tend to have poor conductivity, making them less efficient at conducting heat.

Electrical Conductivity:

Metals are good conductor of electricity, because they have free electrons that can move freely in the metal lattice. Non-metals, with few exception such as graphite, are poor conductor of

electricity because they lack free electrons.

Adaptability:

The metals are malleable and ductile. So, they can be hammered, drawn into wires or transformed into thin sheets without breaking. This property is due to metallic bonds which allow atoms exchange easily under pressure. Non-metals are neither malleable nor ductile rather they are brittle.

Melting Points and Boiling Points

Metals generally have high melting points and boiling points due to strong metallic bonds that require a lot of energy to break. On the other hand non-metals often have lower melting points and boiling points because their atoms and molecules are held by weaker bonds such as covalent bonds, van der Waals bonds, or hydrogen bonds that require less energy to break.

CONCEPT ASSESSMENT EXERCISE 4.10

Compare the general properties of metals and non-metals

KEY POINTS

- When elements are arranged in the order of their increasing atomic number, their properties are repeated in a periodic manner.
- A horizontal row of elements in the periodic table is called a period.
- A column of elements in the periodic table is called a group or a family.
- Group IA and IIA elements are called s-block elements, since s sub-shell fills in these elements.
- Elements in the same group possess similar chemical properties.
- Elements in group IIIA to VIIA are called p-block elements, because filling of valence p sub-shell occurs in these elements.
- The length of a period in the periodic table depends on the type of sub-shell that fills.
- The decrease in force of attraction between nucleus and the valence electron by the electrons present in the inner sub-shells is called shielding effect.
- The size of atom is the average distance between the nucleus of an atom and the outer electronic shell.
- The atomic radii decrease from left to right in a period. Whereas these increase from top to bottom in a group.
- Ionization energy is the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom.
- Electron affinity is 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.

References for additional information

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

REVIEW QUESTIONS

1. Encircle the correct answer.

- Number of periods in the periodic table are:
 - 8
 - 7
 - 16
 - 5
- Which of the following groups contain alkaline earth metals?
 - 1A
 - IIA
 - VIIA
 - VIIIA
- Which of the following elements belongs to VIIIA?
 - Na
 - Mg
 - Br
 - Xe
- Main group elements are arranged in _____ groups.
 - 6
 - 7
 - 8
 - 10
- Period number of ${}_{13}^{27}\text{Al}$ is:
 - 1
 - 2
 - 3
 - 4
- Valence shell electronic configuration of an element M (atomic no. 14) is:
 - $2s^2, 2p^1$
 - $2s^2, 2p^2$
 - $2s^2, 2p^3$
 - $3s^2, 3p^2$
- Which of the following elements you expect to have greater shielding effect?
 - Li
 - Na
 - K
 - Rb
- As you move from right to left across a period, which of the following does not increase:
 - electron affinity
 - ionization energy
 - nuclear charge
 - shielding effect
- All the elements of Group IIA are less reactive than alkali metals. This is because these elements have:
 - high ionization energies
 - relatively greater atomic sizes
 - similar electronic configuration
 - decreased nuclear charge

2. Give short answer.

- Write the valence shell electronic configuration of an element present in the 3rd period and Group IIIA.
- Define halogens.
- Which atom has higher shielding effect, Li or Na?
- Explain why, Na has higher ionization energy than K?
- Alkali metals belong to S-block in the periodic table, why?

3. Arrange the elements in each of the following groups in order of increasing ionization energy:

(a) Li, Na, K

(b) Cl, Br, I

4. Arrange the elements in each of the following in order of decreasing shielding effect.

(a) Li, Na, K

(b) Cl, Br, I

(c) Cl, Br

5. Specify which of the following elements you would expect to have the greatest electron affinity.

S, P, Cl

6. Electronic configuration of some elements are given below, group the elements in pairs that would represent similar chemical properties.

A = $1s^2, 2s^2$ B = $1s^2, 2s^2, 2p^6$ C = $1s^2, 2s^2, 2p^3$ D = $1s^2$ E = $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$ F = $1s^2, 2s^1$ G = $1s^2, 2s^2, 2p^6, 3s^1$ H = $1s^2, 2s^2, 2p^6, 3s^2$

7. Arrange the elements in groups and periods in Q. No. 6.

IA								VIIIA	
	IIA								
			IIIA	IVA	VA	VIA	VIIA		

8. For normal elements, the number of valence electrons of an element is equal to the group number. Find the group number of the following elements.

 ${}_{13}^{27}\text{Al}$, ${}_{16}^{32}\text{S}$, ${}_{19}^{39}\text{K}$, ${}_{6}^{8}\text{O}$

Atomic number	Mass number	No. of protons	No. of neutrons	No. of electrons
11			12	
		14	15	
	47		25	
	27			13

- 13 In which block, group and period in the periodic table would you place each of the following elements with the following electronic configurations?
- (a) $1s^2, 2s^1$ (b) $1s^2, 2s^2, 2p^5$
(c) $1s^2, 2s^2, 2p^6, 3s^2$ (d) $1s^2$

14. What types of elements have the highest ionization energies and what types of elements have the lowest ionization energies. Argue.

15. i. Two atoms have electronic configuration $1s^2, 2s^2, 2p^6$ and $1s^2, 2s^2, 2p^6, 3s^1$. The ionization energy of one is 2080 kJ/mole and that of the other is 496 kJ/mole. Match each ionization energy with one of the given electronic configuration. Give reason for your choice.

ii. Use the second member of each group from Group IA, IIA and VIIA to judge that the number of valence electron in an atom of the element is the same as its group number.

iii. Letter A, B, C, D, E, F indicates elements in the following figure:

					C		
A			B				
	D			E			
					F		

- a. Which elements are in the same periods?
 - b. Write valence shell electronic configuration of element D.
 - c. Which elements are metals?
 - d. Which element can lose two electrons?
 - e. In which group E is present?
 - f. Which of the element is halogen?
 - g. Which element will form dipositive cation?
 - h. Write electronic configuration of element E
 - i. Which two elements can form ionic bond?
 - j. Can element C form C_2 molecule? Interpret.
 - k. Which element can form covalent bonds?
 - l. Is element F a metal or non-metal?
16. Electronic configurations of four elements are given below:
 (a) $1s^2, 2s^1$ (b) $1s^2, 2s^2, 2p^5$ (c) $1s^2, 2s^2, 2p^6, 3s^2$ (d) $1s^2$
 Which of these elements is
- i) An alkali metal
 - ii) An alkaline earth metal
 - iii) A noble gas
 - iv) A halogen
17. Argue in what region of the periodic table you will find elements with relatively
- a) high ionization energies
 - b) low ionization energies

PROJECT ←

Prepare 3D model of the periodic table (Group Activity)

