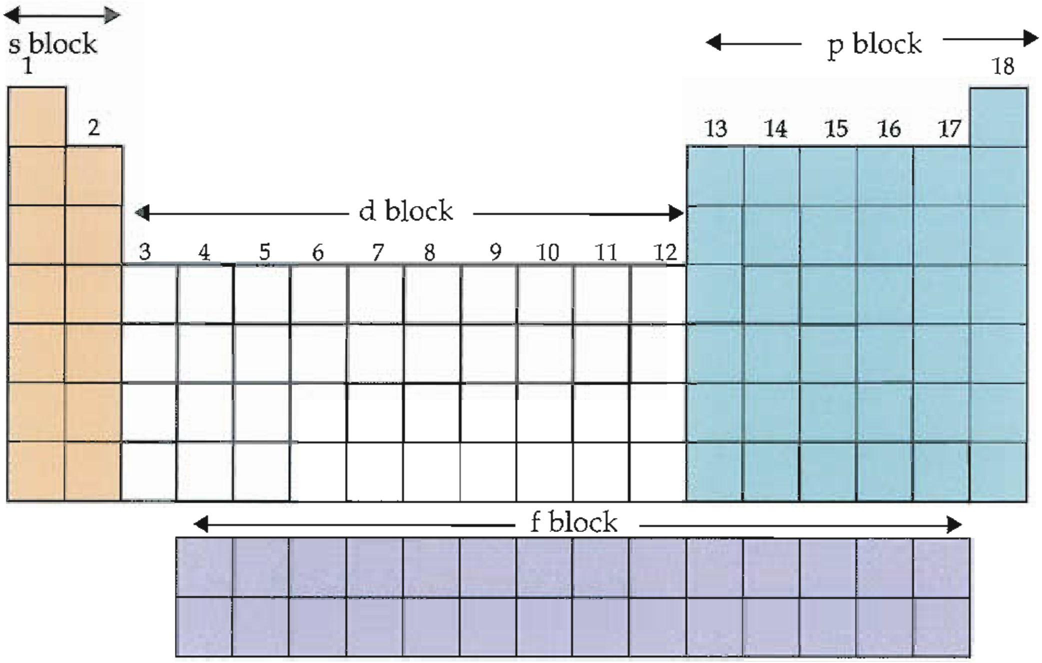
3

ELECTRONIC CONFIGURATION AND THE PERIODIC TABLE



Think of an element x. Say its atomic number is 12.

Can you predict its position in the periodic table?

Electronic configuration - 2, 8, 2

Group number - ....

Period number - ....

* Is the element a metal or a non-metal?

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* How did you know that?

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* Let the position of another element y be just above that of x in the periodic table (both symbols not real). What will the electronic configuration of y be?

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* Are the valencies of both x and y the same? How could you infer this?

Isn't it your knowledge of the relationship between elements in the periodic table and their electronic configuration that helped you to predict this? This points out the importance of the periodic table in the study of elements.

* See the periodic table given above. Do you observe that the elements are classified into 4 blocks? What are they?

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* Which are the blocks containing representative elements?

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* By what name were the d-block elements familiar to you?

Is there any basis for classifying elements in the periodic table into blocks? Let us examine.

You have understood that the basic reason for the periodicity in the properties of elements in the periodic table is their electronic configuration. The properties of elements are influenced mainly by the electronic arrangement in their atoms. You are familiar with the Bohr scheme of arrangement of electrons in an atom. According to this, electrons that move around the nucleus have definite paths.

What is the name of such a fixed path?

Electrons present in each such shell or orbit have a definite energy. Therefore shells are also known as energy levels.

You know that the Bohr model of atom has shells K, L, M, N ...... These can also be referred to as 1, 2, 3, 4 ...... energy levels respectively (Fig. 3.1). Try to write down the maximum number of electrons that can be accommodated in each shell:

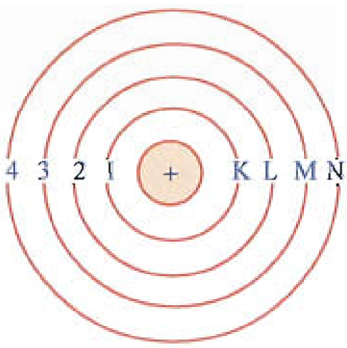

Fig. 3.1

K -

L -

M -

N -

An electron in the K shell has a fixed energy. Suppose you want to transfer an electron from the K shell to the L shell. It is found that a definite amount of energy has to be supplied for this. If so, the energy of electrons in L shell should be greater than that of electrons in K shell. That means the energy of L shell is higher than that of K shell.

* Is the energy of electrons in the M shell higher or lower than that of electrons in the K and L shells?

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* What happens to the energy of shells when their distance from the nucleus increases?

  Decreases / increases (✓ the correct choice).

Electrons are accommodated in shells in the increasing order of energy. That is, electrons are filled first in the lowest energy orbit and then in the next higher energy orbit and so on.

* If an atom has only one electron, in which shell will it be filled?

----

* And if there are two electrons?

----

* Where does the third electron go? To the L or M shell?

Now, you might have understood that the shells in an atom are filled with electrons in the increasing order of their energy.

Subshells

The Bohr model provides a simple explanation of the electronic structure of an atom. But the Bohr model became inadequate to explain the properties of elements of higher atomic numbers. Studies related to the position and properties of electrons in an atom led to the conclusion that the electrons in each energy level are distributed in the various sub energy levels in it. These sub energy levels are called subshells. They are named as s, p, d, f.... . The number of subshells in a shell will be equal to the serial number of the shell. That means the first shell has only one subshell (s), the second has two (s, p) and so on. Each subshell can accommodate a fixed number of electrons. The s subshell can accommodate a maximum of 2 electrons. The K shell has only the s subshell and hence it can have only 2 electrons. Go through Table 3.1 which shows various shells, their subshells and the maximum number of electrons that can be accommodated in each of them. Try to complete the table.

- ★ Have you noticed from Table 3.1 that the same subshell may occur in different shells?

  Name the subshell common to all shells.

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- ★ To identify the shell to which the subshell belongs, the subshells are represented by including the shell number also as 1s, 2s, 2p..... etc.

Let us see the maximum number of electrons that can be filled in each subshell. See how the s subshell of the K shell is filled. It has only two electrons. Out of the 8 electrons in the L shell, two are in the s subshell. Now what is the maximum number of electrons in the p subshell? Fill this in the gap above the name of the subshell against the L shell.

We have seen the maximum number of electrons in s and p sub shells. Now, find the maximum number of electrons that can be accommodated in the d subshell by filling the gaps against the third shell in the table. Can't you find the maximum number of electrons in the f subshell by filling the gaps against the fourth shell in the table in a similar way?

Main energy levels		Maximum number of electrons accommodated	Number of electrons accommodated and name of the subshell				Number of subshells
Sl. No. of shell	shell						
1	K	2	2 1s				1
2	L	8	2 2s	--- 2p			2
3	M	18	--- --s	--- --p	--- --d		----
4	N	32	--- --s	--- --p	--- --d	--- --f	----

Table 3.1

Consolidate your findings in Table 3.2.

Subshells	s	p	d	f
Maximum number of electrons accommodated	---	---	---	---

Table 3.2

You have understood that each shell has a definite energy. But intensive studies have proved that subshells in a shell differ slightly in their energy. Because of this, the energy of electrons present in a shell may vary slightly. The increasing order of energy of subshells is represented as

s < p < d < f ....

Try to write the subshells in the fourth shell in the increasing order of their energy.

4s < --- < --- < ---

Subshell electronic configuration

Aren't electrons accommodated in shells in the increasing order of their energy? When electrons are filled in a shell of an atom they are actually accommodated in the respective subshells of the shell. Let us see how subshell electronic configuration of elements is written.

You know that the number of electrons in an atom is equal to its atomic number (Z).

Atomic number of hydrogen is 1.

- Then, how many electrons?

- In which shell?

- In which subshell?

The subshell electronic configuration of hydrogen can be represented as

$$1s^1$$

(It is read as one s one. The number at the right top shows the number of electrons).

What about helium?

Atomic number - 2

Number of electrons - ....

Subshell filled - s

Number of electrons in s subshell - 2

Subshell electronic configuration - $1s^2$

- Number of electrons in the atom of Li (Z = 3)

- Try to write the subshell electronic configuration - $1s^2 2s^{....}$
- Be (Z =4) - $1s^{....} 2s^{....}$

The next element is boron, isn't it?

Atomic number –

Number of electrons –

Subshell electronic configuration –

$1s^{....} 2s^{....} 2p^{....}$

Now try to write the subshell electronic configuration of $_6C$, $_8O$, $_9F$, $_{10}Ne$, $_{11}Na$..... also like this. Identify in the periodic table, the blocks to which these elements belong. Fill the data in Table 3.3.

Do you find any peculiarity in the subshell electronic configuration of elements in s and p blocks? Examine the relationship between the block in which the element is included and the subshell to which the last electron of these elements gets filled.

MODERN PERIODIC TABLE

s block elements

p block elements

d block elements

Atomic number Name Symbol Atomic mass Electronic Configuration

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 HYDROGEN **H** 1.008 1																	2 HELIUM **He** 4.0 2
3 LITHIUM **Li** 6.9 2.1	4 BERYLLIUM **Be** 9.0 2.2											5 BORON **B** 10.8 2.3	6 CARBON **C** 12.0 2.4	7 NITROGEN **N** 14.0 2.5	8 OXYGEN **O** 16.0 2.6	9 FLUORINE **F** 19.0 2.7	10 NEON **Ne** 20.2 2.8
11 SODIUM **Na** 23.0 2.8.1	12 MAGNESIUM **Mg** 24.3 2.8.2											13 ALUMINIUM **Al** 27.0 2.8.3	14 SILICON **Si** 28.1 2.8.4	15 PHOSPHORUS **P** 31.0 2.8.5	16 SULPHUR **S** 32.1 2.8.6	17 CHLORINE **Cl** 35.5 2.8.7	18 ARGON **Ar** 39.9 2.8.8
19 POTASSIUM **K** 39.1 2.8.8.1	20 CALCIUM **Ca** 40.1 2.8.8.2	21 SCANDIUM **Sc** 45.0 2.8.9.2	22 TITANIUM **Ti** 47.9 2.8.10.2	23 VANADIUM **V** 50.9 2.8.11.2	24 CHROMIUM **Cr** 52.0 2.8.13.1	25 MANGANESE **Mn** 54.9 2.8.13.2	26 IRON **Fe** 55.8 2.8.14.2	27 COBALT **Co** 58.9 2.8.15.2	28 NICKEL **Ni** 58.7 2.8.16.2	29 COPPER **Cu** 63.5 2.8.18.1	30 ZINC **Zn** 65.4 2.8.18.2	31 GALLIUM **Ga** 69.7 2.8.18.3	32 GERMANIUM **Ge** 72.6 2.8.18.4	33 ARSENIC **As** 74.9 2.8.18.5	34 SELENIUM **Se** 79.0 2.8.18.6	35 BROMINE **Br** 79.5 2.8.18.7	36 KRYPTON **Kr** 83.8 2.8.18.8
37 RUBIDIUM **Rb** 85.5 2.8.18.8.1	38 STRONTIUM **Sr** 87.6 2.8.18.8.2	39 YTTRIUM **Y** 88.9 2.8.18.9.2	40 ZIRCONIUM **Zr** 91.2 2.8.18.10.2	41 NIOBIUM **Nb** 92.9 2.8.18.12.1	42 MOLYBDENUM **Mo** 95.9 2.8.18.13.1	43 TECHNETIUM **Tc** (99) 2.8.18.14.1	44 RUTHENIUM **Ru** 101.1 2.8.18.15.1	45 RHODIUM **Rh** 102.9 2.8.18.16.1	46 PALLADIUM **Pd** 106.4 2.8.18.18	47 SILVER **Ag** 107.9 2.8.18.18.1	48 CADMIUM **Cd** 112.4 2.8.18.18.2	49 INDIUM **In** 114.8 2.8.18.18.3	50 TIN **Sn** 118.7 2.8.18.18.4	51 ANTIMONY **Sb** 121.8 2.8.18.18.5	52 TELLURIUM **Te** 127.6 2.8.18.18.6	53 IODINE **I** 126.9 2.8.18.18.7	54 XENON **Xe** 131.3 2.8.18.18.8
55 CAESIUM **Cs** 132.9 2.8.18.18.8.1	56 BARIUM **Ba** 137.3 2.8.18.18.8.2	57 LANTHANUM **La** 138.9 2.8.18.18.9.2	72 HAFNIUM **Hf** 178.5 2.8.18.32.10.2	73 TANTALUM **Ta** 181.0 2.8.18.32.11.2	74 TUNGSTEN **W** 183.9 2.8.18.32.12.2	75 RHENIUM **Re** 186.2 2.8.18.32.13.2	76 OSMIUM **Os** 190.2 2.8.18.32.14.2	77 IRIDIUM **Ir** 192.2 2.8.18.32.15.2	78 PLATINUM **Pt** 195.1 2.8.18.32.17.1	79 GOLD **Au** 197.0 2.8.18.32.18.1	80 MERCURY **Hg** 200.6 2.8.18.32.18.2	81 THALLIUM **Tl** 204.4 2.8.18.32.18.3	82 LEAD **Pb** 207.2 2.8.18.32.18.4	83 BISMUTH **Bi** 209.0 2.8.18.32.18.5	84 POLONIUM **Po** (209) 2.8.18.32.18.6	85 ASTATINE **At** (210) 2.8.18.32.18.7	86 RADON **Rn** (222) 2.8.18.32.18.8
87 FRANCIUM **Fr** (223) 2.8.18.32.18.8.1	88 RADIUM **Ra** (226) 2.8.18.32.18.8.2	89 ACTINIUM **Ac** (227) 2.8.18.32.18.9.2	104 RUTHERFORDIUM **Rf** (260)	105 DUBNIUM **Db**	106 SEABORGIUM **Sg**	107 BOHRIUM **Bh**	108 HASSIUM **Hs**	109 MEITNERIUM **Mt**	110 DARMSTADTIUM **Ds**	111 ROENTGENIUM **Rg**	112 UNUNBIUM **Uub**	113 UNUNTRIUM **Uut**	114 UNUNQUADIUM **Uuq**	115 UNUNPENTIUM **Uup**	116 UNUNHEXIUM **Uuh''''**		

f block elements

Lanthanones	58 CERIUM **Ce** 140.1 2.8.18.20.8.2	59 PRASEODYMIUM **Pr** 140.9 2.8.18.21.8.2	60 NEODYMIUM **Nd** 144.2 2.8.18.22.8.2	61 PROMETHIUM **Pm** 145 2.8.18.23.8.2	62 SAMARIUM **Sm** 150.4 2.8.18.24.8.2	63 EUROPIUM **Eu** 152.0 2.8.18.25.8.2	64 GADOLINIUM **Gd** 157.3 2.8.18.25.9.2	65 TERBIUM **Tb** 158.9 2.8.18.27.8.2	66 DYSPROSIUM **Dy** 162.5 2.8.18.28.8.2	67 HOLMIUM **Ho** 165 2.8.18.29.8.2	68 ERBIUM **Er** 167.3 2.8.18.30.8.2	69 THULIUM **Tm** 168.9 2.8.18.31.8.2	70 YTTERBIUM **Yb** 173.0 2.8.18.32.8.2	71 LUTETIUM **Lu** 175.0 2.8.18.32.9.2
Actinones	90 THORIUM **Th** 232 2.8.18.32.18.10.2	91 PROTACTINIUM **Pa** 231 2.8.18.32.20.9.2	92 URANIUM **U** 238.0 2.8.18.32.21.9.2	93 NEPTUNIUM **Np** (237) 2.8.18.32.22.9.2	94 PLUTONIUM **Pu** (244) 2.8.18.32.23.9.2	95 AMERICIUM **Am** (243) 2.8.18.32.24.9.2	96 CURIUM **Cm** (247) 2.8.18.32.25.9.2	97 BERKELIUM **Bk** (247) 2.8.18.32.26.9.2	98 CALIFORNIUM **Cf** (251) 2.8.18.32.27.9.2	99 EINSTEINIUM **Es** (252) 2.8.18.32.28.9.2	100 FERMIUM **Fm** (257) 2.8.18.32.29.9.2	101 MENDELEVIUM **Md** (258) 2.8.18.32.30.9.2	102 NOBELIUM **No** (259) 2.8.18.32.31.9.2	103 LAWRENCIUM **Lr** (260) 2.8.18.32.32.9.2

The element with atomic number 112 has been named as coppernitium. Symbol Cn.

s block		p block	
Element	Subshell Electronic configuration	Element	Subshell Electronic configuration

Table 3.3

Our observation.

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* Can't Table 3.3 be expanded by including more elements?

Helium (He)

The outermost electron in helium fills in the s subshell. Therefore it should have been placed in the s block. But helium being a noble gas is placed in the p block along with other noble gases.

Write down the electronic configuration of Cl (Z=17) and find the block to which it belongs. Is the position of chlorine in the same block that you identified? Verify your answer referring the periodic table.

What about Ar (Z=18)?

* Its subshell electronic configuration is

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* Block

Atomic number of potassium is 19. What will its subshell electronic configuration be? Look at the answer given by a student

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1$

Is it correct? According to this configuration to which block does potassium belong?

Does potassium belong to d block? Verify using the periodic table.

* What did you find?

Let us examine the reason for including potassium in s block. Potassium is an element in the fourth period of the periodic table. Its shell-wise electronic configuration is 2, 8, 8, 1. How many shells are present?

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* In which shell does the last electron enter? In which subshell?

----

* Which are the subshells in the fourth energy level?

  4s, ..., ..., ...,

* In which subshell does the electron fill first?

----

* In potassium the electron goes to 4s subshell in preference to 3d. Isn't this the reason for potassium being an s block element?

- ★ If so, between 4s and 3d, which has the lower energy?

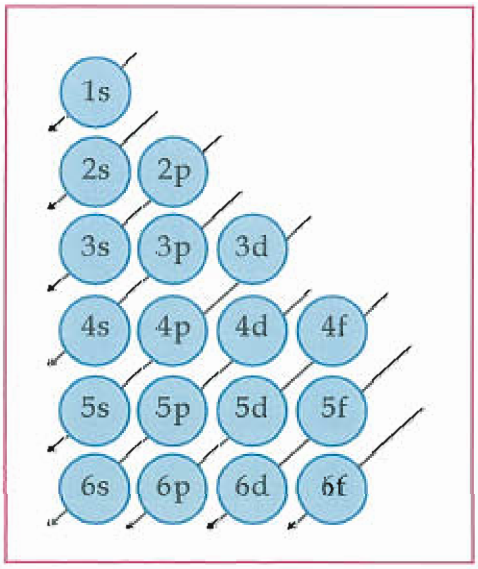
- ★ Now try to write down the correct subshell electronic configuration of potassium considering this too.

- ★ Isn't the correct subshell electronic configuration of potassium $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$? Confirm this by comparing with the shell - wise configuration.

Examine whether this is true in the case of other elements also that follow potassium. Write down how electrons are filled in various subshells in the increasing order of their energy.

1s < 2s < 2p < ...... < ...... < ...... < ...... < ......

The following diagram may be used for this.



Now write down the electronic configuration of elements $_{20}Ca$, $_{21}Sc$, $_{22}Ti$ and $_{23}V$. Do they agree with the positions of the elements in the periodic table?

See how the electronic configuration of scandium (Z=21) is to be written:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$

- ★ Though electrons are filled in 3d subshell only after 4s, the subshells of the same shell are to be put together while writing the electronic configuration.
- ★ Like this how will you write the electronic configuration of chromium (Z=24)?

- ★ Aren't there only four electrons in the 3d subshell?

But in the case of chromium instead of the 2 electrons filling in the outermost 4s subshell one of them goes to 3d subshell. Its configuration, therefore, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$.

The electronic configuration of copper (Z=29) also has such a rearrangement. Its configuration is

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

The d^5 and d^{10} are stable electronic configurations in the d subshell. The rearrangement of electrons in Cr and Cu is to attain these stable electronic configurations.

- ★ Write the subshell electronic configuration of the elements of first four periods in the periodic table in the order of their atomic numbers (1 to 36) as a table in your school diary.

Elements of s and p blocks

Haven't you understood the basis of classification of elements into the four blocks in the periodic table from their electronic configurations written?

Group number / Period number	1 (s block)	2 (s block)	13 (p block)	14 (p block)	15 (p block)	16 (p block)	17 (p block)	18 (p block)
2	3 Li $2s^1$	4 Be $2s^2$	5 B $2s^22p^1$	6 C $2s^22p^2$	7 N $2s^22p^3$	8 O $2s^22p^4$	9 F $2s^22p^5$	10 Ne $2s^22p^6$
3	11 Na $3s^1$	12 Mg $3s^2$	13 Al $3s^23p^1$	14 Si $3s^23p^2$	15 P $3s^23p^3$	16 S $3s^23p^4$	17 Cl $3s^23p^5$	18 Ar $3s^23p^6$

Table 3.4

Which are the groups coming in the s and p blocks?

- ★ s block -- group 1 and group 2
- ★ p block --

\- -

A part of the periodic table containing s and p block elements along with their subshell-wise electronic configurations of the outermost shell is given. (Table 3.4)

- ★ In which subshell does the last electron of each element of groups 1 and 2 fall?

\- -

- ★ How many electrons are present in the s subshell of the outermost shell in the elements of group 1?

\- -

- ★ And in group 2 elements?

\- -

- ★ Thus in the case of s block elements, isn't it possible to find the relationship between the group number of the element and the number of electrons in its outermost subshell.
- ★ Examine also the relationship between the serial number of the outermost filled shell of an element and period in which it is present.
- ★ Our finding

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Now consider the elements in p block.

The subshell-wise electronic configuration of the outermost shell of boron is $2s^2\ 2p^1$. Find its position in the periodic table using Table 3.4.

The outermost subshell number.............

Period number ....

Total number of electrons in the subshells of the outermost shell ....

Group number ....

In the same way find out the electronic configuration of other p block elements such as carbon, oxygen, neon, aluminium etc. and fix their positions.

Can you suggest a method to find out the group number of an element from the total number of electrons in the subshells of its outermost shell?

\- -

\- -

Don't you see that the outermost shell number and period number of an element are the same?

The subshell electronic configuration of an element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$. The outermost shell configuration is $4s^24p^3$.

Now try to predict its period, group and block.

Gallium is the element that comes just below aluminium in group 13. What will be the subshell configuration of its outermost shell?

Now you can write this without referring the periodic table, can't you? Isn't this due to the similarity in the subshell electronic configuration after regular intervals?

Isn't this the reason for similarities in the properties of elements in a group also?

Now let us see what all characteristics of elements can be predicted by analysing the subshell electronic configuration of elements in this way.

Element	Outermost electronic configuration	Complete electronic configuration	Atomic number	Period	Group	Block
X	$3s^2$					
Y	$3s^2 3p^5$					

Table 3.5

Fill up the blanks in Table 3.5

(symbols X and Y are not real)

- ✶ Which is the element with valency 1?

- ✶ Write down the formula of the possible compound formed by X and Y.

- ✶ Which is the element that shows metallic character?

Elements of d block

You have already made out that d block elements are transition elements. We also know that all of them are metals.

List the familiar ones.

Elements from group 3 to 12 belong to d block. Now see the electronic arrangement of the 3d and 4s, the last subshells of the fourth period elements given in Table 3.6

Group →

3	4	5	6	7	8	9	10	11	12
21 Sc $3d^1 4s^2$	22 Ti $3d^2 4s^2$	23 V $3d^3 4s^2$	24 Cr $3d^5 4s^1$	25 Mn $3d^5 4s^2$	26 Fe $3d^6 4s^2$	27 Co $3d^7 4s^2$	28 Ni $3d^8 4s^2$	29 Cu $3d^{10} 4s^1$	30 Zn $3d^{10} 4s^2$

Table 3.6

★ Which is the outermost subshell in these elements? What is the common feature in the number of electrons filled in this subshell (except for Cr and Cu)?

★ To which shell and subshell do the electrons go?

★ Is there any relationship between the total number of electrons in 3d and 4s subshells and the group number?

Find out this relationship from Table 3.6

In s and p block elements the last electrons enter the outermost shell. Is this the same for d block elements?

Similarly s and p block elements show similarities in properties in their groups. But transition elements show similarity in a group as well as along a period. What might the reason be?

- The properties of elements are decided mainly by the electronic configuration in the outermost shell of their atoms. The s subshell of the outermost shell of most of the transition elements in a period has the same number of electrons. Isn't this the reason for similarities in their properties?

With the help of the information given above prepare a note comparing the special features of the electronic configuration of s, p and d block elements and present it in your class.

What are the peculiarities of d block elements that differ from s and p block elements?

Look at the names and formulae of two compounds formed by iron with chlorine

- ferrous chloride $FeCl_2$
- ferric chloride $FeCl_3$

You are already familiar with writing the formulae of compounds by inter changing the valencies to suffix the symbols of elements. If so, try to find the valencies of elements in the above compounds.

Compound	Valency	
	Fe	Cl
$FeCl_2$		
$FeCl_3$		

Table 3.7

★ Which element shows variable valency?

★ Does the metallic element in a compound generally donate or accept electrons?

★ How many electrons are released by Fe in $FeCl_2$?

★ And in $FeCl_3$?

Write down the symbols indicating the charge on each:

Ferrous ion Fe^{2+}

Ferric ion ....

The subshell configuration of iron ($_{26}Fe$) is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$. Then what will be the electronic configuration of Fe^{2+} and Fe^{3+}?

Fe^{2+} -----------------

Fe^{3+} -----------------

From which subshell are electrons lost to form Fe^{2+}?

Are electrons lost only from the outermost shell while forming Fe^{3+} ion?

The energy difference between the s subshell of the outermost shell and the d subshell of the penultimate shell in d block elements is very small. Therefore during compound formation electrons from both these subshells may be transferred or shared. Hence these elements can show variable valencies. Examine whether this is true in the case of other elements of d block also.

Copper forms two chlorides CuCl and $CuCl_2$. Is the valency of Cu the same in both the compounds? Try to write.

What is the electronic configuration of $_{29}Cu$?

- - - - - - - - - - - - - - - -

Find out the subshells from which electrons are lost, by writing the electronic configuration of Cu^+ and Cu^{2+}

Cu^+ - - - - - - - - - - - - - - - -

Cu^{2+} - - - - - - - - - - - - - - - -

Aren't you convinced that d block elements show variable valency?

See another important feature of d block elements.

Prepare aqueous solutions of ferrous sulphate and ferric sulphate in two test tubes. Add equal quantities of sodium hydroxide solution to both. What change did you observe? Enter them in Table 3.8.

Aren't the colours obtained in the two different? What are the compounds responsible for the colour?

- -

Collect the samples of compounds listed below from your school laboratory and examine their colours.

Compound	colour
Sodium sulphate	Colourless
Copper sulphate	________
Ferrous sulphate	________
Magnesium sulphate	________
Cupric chloride	________
Cobalt nitrate	________
Sodium chloride	________
Calcium carbonate	________
Ammonium dichromate	________
Potassium nitrate	________
Manganese dioxide	________

* List the metals in the coloured compounds given above.

- -

* Find the block to which these metals belong with the help of the periodic table.

Transition metal compounds are used to give colour to glasses and oil paintings. Aren't their different colours the reason for this?

Can't you now make a note consolidating the characteristic features of d block

	Colour change	Products formed
$FeSO_4$ + NaOH	....	$Fe(OH)_2 + Na_2SO_4$
$Fe_2(SO_4)_3$ + NaOH	....	$Fe(OH)_3 + Na_2SO_4$

Table 3.8

elements on the basis of the above discussion?

Elements of f block

The f block elements include lanthanones and actinones. The last electron in the atoms of these elements is filled in the f subshell. These elements are presented as a separate block in the periodic table though they are part of the sixth and seventh periods. Just imagine how big the periodic table will be if they were not presented so! Will it be compact? Lanthanones of the f block are commonly known as rare earths. They are found in the beach sands of Kerala. Uranium and thorium are important among actinones. Most of the actinones are man-made and radioactive elements.

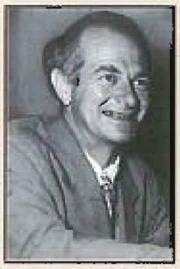
Periodic trends in the properties of elements

The main reason for periodicity in the properties of elements is their electronic configuration. Properties of elements such as valency, electronegative and electropositive nature, atomic size etc., are related to their electronic configuration. Properties will also change when there is a change in the electronic configuration.

Electronegativity

Electronegativity is the capacity of each atom involved in bonding to attract bonded electrons towards it. You know that non-metals have comparatively high electronegativity. Consider a covalent bond formed between two atoms A and B by sharing a pair of electrons. Since the capacities of A and B to attract the bonded electrons differ, the electrons move in the direction of greater attraction. Suppose it is towards B. Then, B is said to be more electronegative than A.

Pauling Scale



The Pauling scale is used to compare the electronegativity of various elements. It was developed by the American scientist Linus Pauling. In this scale electronegativity values of elements vary from zero to four.

See a portion of the periodic table containing electronegativities of elements given in Table 3.9. Find the element with the highest electronegativity from the table. Is it a metal or non-metal?

What change occurs in the electronegativities of elements while

H 2.1						
Li 1	Be 1.5	B 2	C 2.5	N 3	O 3.5	F 4
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 08.						Br 2.8
Rb 0.8						I 2.5
Cs 0.7						At 2.2

Table 3.9

moving from left to right along a period?

increases / decreases (✓ the correct choice).

And on moving down a group?

Consider the bonding between carbon and chlorine.

Name the compound formed between them?

Write its formula.

What type of bond is formed between the elements in this compound?

ionic / covalent (✓ the correct choice)

Electronegativity values of carbon and chlorine are given below. Find the difference between them.

C - 2.5, Cl - 3.00

difference = 3.00 - 2.5 = ...

Electronegativity values of sodium and chlorine are 0.90 and 3.00. Can't we calculate the difference

3.00 - 0.9 = ...

You know that the bond in sodium chloride is ionic.

When the electronegativity difference between the elements increases there is the possibility for ionic bonding and when it decreases, for covalent bonding. If the difference is above 1.7 there is the chance for the formation of an ionic compound. Using the electronegativity values of elements, find out the type of bonds present in the following compounds:

$MgCl_2$, Na_2O, CCl_4, HCl

In the diatomic molecules of Cl_2, O_2, H_2 and F_2 both the atoms in each have the same electronegativity value. They attract the bonded electrons equally. Hence the bond is considered as a pure covalent bond.

See the representation of HCl molecule. Here also the bond is covalent.

★ Which of the atoms is more electronegative?

★ Which atom will attract the bonded electron pair more towards it?

★ From which of the atoms does the electron move away?

If the electronegativity values of the two atoms involved in covalent bonding are different, the bonded electrons show a tendency to move closer towards the atom of high electronegativity. As a result the atom of higher electronegativity gets a partial negative charge and the other with lower electronegativity gets a partial positive charge. These type of bonds are said to be polar covalent bonds. The partial charge on atoms can be represented as δ^+ and δ^-.

Which atom of HCl molecule gets δ^+ charge?

And which gets δ^- charge?

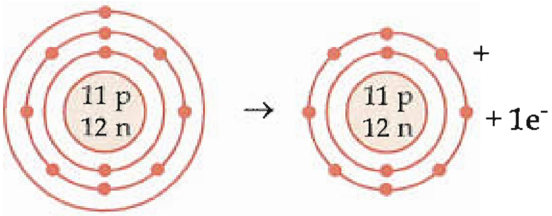
The polar nature of HCl molecule can be represented by

$H^{\delta+} - Cl^{\delta-}$

These type of molecules are generally called polar molecules. They have partial ionic character. H_2O, HF, NH_3 etc. are polar molecules.

Ionisation energy

See the illustration of an isolated atom changing into an ion.



Sodium atom (Na) → Sodium ion (Na^+)+1e⁻

You know that the nucleus of an atom attracts electrons. Shouldn't energy be supplied to overcome this force of attraction and remove the electron from the sodium atom?

The minimum energy required to remove the most loosely bound electron from an isolated gaseous atom is called its ionisation energy. It is expressed in kJ / mol

The position of potassium is just below sodium among the first group elements. What is the change in atomic size while moving from top to bottom in a group?

increases / decreases (✓ the correct choice)

- ★ Which atom is bigger-sodium or potassium?

  -

- ★ In which among these atoms is the attraction of nucleus towards the outermost electrons more?

  -

- ★ Suppose the electron in each atom is to be removed. Which may need higher ionisation energy. Why?

  -

  -

- ★ Now try to write the relationship between atomic size and ionisation energy.

  With increase in atomic size ionisation energy

  increases / decreases (✓ the correct choice)

- ★ And when atomic size decreases?

  -

Let us now examine the variation of ionisation energy of elements in relation to their atomic size in a group and a period.

As we move down a group

- ★ atomic size

  -

- ★ ionisation energy

  -

When we move from left to right along a period

- ★ atomic size

  -

- ★ ionisation energy

  -

- ★ When we compare the ionisation energy of metals and non-metals in a period, which of them have lower ionisation energy? Why?

  -

Metals generally lose electrons in a chemical reaction. So they have electropositive nature. Isn't their low ionisation energy responsible for their electropositive nature?

You have already learnt through experiment that potassium is more reactive than sodium. Analyse the reason for this by relating to their ionisation energies. Where are the elements of high ionisation energy placed in the periodic table?

On the left side or right side?

Aren't they generally non-metals?

The elements with the highest ionisation energy are the noble gases in group 18.

Examine the ionisation energy of a few elements given below (symbols not real). One of them is that of an inert gas.

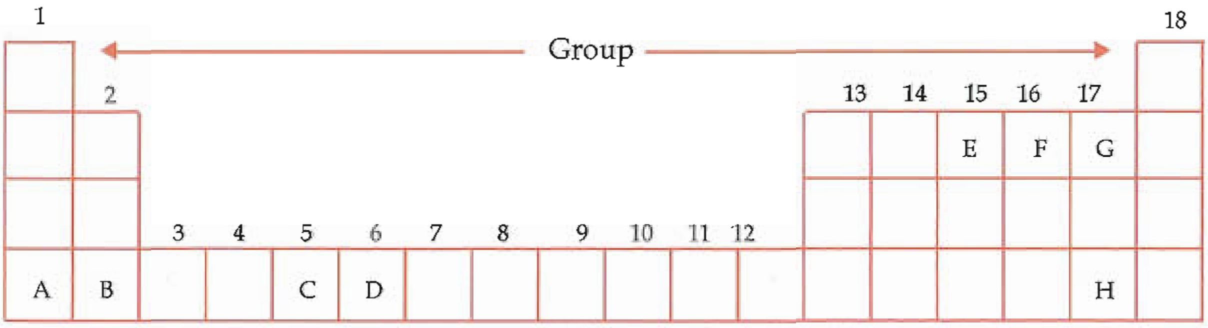


Table 3.10

A - 2372 kJ/mol

B - 520 kJ/mol

C - 900 kJ/mol

D - 1680 kJ/mol

★ Which is the metal with high reactivity? Why?

\- -

★ Which of them is likely to be an inert gas?

\- -

★ Arrange the elements from left to right as in the periodic table.

See the portion of the periodic table given in Table 3.10.

Symbols of elements given are not real.

- Which of the elements has 5 electrons in the outermost shell?
- Which elements have 5 electrons in the outermost subshell?
- Which element has the configuration $3d^3\ 4s^2$ in the two outer subshells?
- Which element can accept two electrons to form ionic bond?
- Which element has the highest ionisation energy?
- Which element has the lowest ionisation energy?
- Which is the most reactive metal?
- Which is the most reactive non-metal?
- Which element has the lowest electronegativity?
-
-

Find more questions of this kind.

Thus, in this chapter, we have come across the possibilities of analysing and comparing the characteristics of elements using the periodic table in the study of chemistry. From the discussion so far, aren't you convinced how much useful the periodic table is to simplify and systematize the study of the elements which are the basic units of all substances and also about the millions of compounds formed by their combination?

Follow up Activities

1. A few subshells of an atom are given below.

   4d, 5s, 3f, 2p

   (a) Which subshell is not possible?

   (b) What is the reason for this?

2. Look at the subshell electronic configuration of some atoms given below.

   $1s^2\ 2s^2\ 3s^2$

   $1s^2\ 2s^2\ 2p^6\ 3s^1$

   $1s^2\ 2s^2\ 2p^7$

   $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$

   $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

   (a) Which are the wrong configurations?

   (b) Give the reason.

3. Analyse the table given below and answer the following questions. (symbols are not real)

Element	A	B	C	D
Mass number	4	23	40	16
Number of neutrons	2	12	22	8

   (a) What is the atomic number of element B?

   (b) Write down the subshell electronic configuration of element D?

   (c) Which are the noble elements? Why?

   (d) Arrange the elements in the increasing order of their ionisation energy.

4. The positions of elements A and B in the periodic table are given below (symbol not real).

   A - period - 2, group - 17

   B - period - 3 , group - 2

   (a) Write down the subshell electronic configurations of A and B.

   (b) What is the chemical formula of the compound formed between A and B? What type of bond will it have?

5. A few characteristics of the elements in group 17 are given in the following table. Complete the table by filling the blanks and answer the following questions.

Element	Symbol	Physical state at STP	Reactivity with hydrogen	Formula of the hydride
Fluorine	F	Gas	Vigorous reaction	HF
....	Cl	....	Vigorous reaction	....
Bromine	....	Liquid	Slow reaction	....
Iodine	I	Solid	Very slow reaction	HI

(a) What is the name by which the family of group 17 elements known?

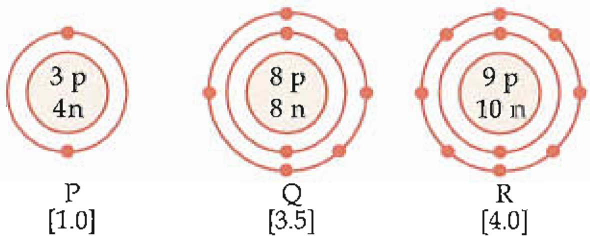
(b) What is their common valency?

(c) Which element has the highest electronegativity?

(d) What is the trend in the reactivity of these elements from top to bottom in the group? increases / decreases

(e) How is the reactivity of the non-metals related to their electronegativity?

6. Examine the Bohr models of atoms/ions of elements P, Q and R given (symbols not real). The quantities given in brackets are the electronegativity values of the elements.



(a) Which are the models representing the ions?

(b) Which is likely to be a positive ion? Why?

(c) Write the symbol of the negative ion.

(d) Which element is likely to be seen at the left side of the periodic table?

(e) Assume that the elements P and Q enter into a chemical reaction.

(i) Write the balanced equation of the reaction.

(ii) What would be the type of bond formed between them?

(f) Which is the most reactive non-metal? Why?