3

PERIODIC TABLE AND PERIODICITY OF PROPERTIES

This is a 20 days lesson

After completing this lesson, you will be able to:

- Classify the element (into two categories: groups and periods) according to the configuration of their outer most shell.
- Distinguish between a period and a group in the periodic table. (Understanding)
- State the periodic law.
- Determine the demarcation of the periodic table into an s block and p block.
- Determine the location of families in the periodic table. Recognize the similarity in the chemical and physical properties of elements in the same family of elements.
- Describe how electronegativites change within a group and within a period in the periodic table.
- Explain the shape of the periodic table.
- Explain how shielding effect influences periodic trends.
- Identify the relationship between electronic configuration and the position of an element in the periodic table.

Pre- Reading

By the end of 18th century, 23 elements were known, by 1870, 65, by 1925, 88, today there are 109. These elements combine to form millions of compounds. It is very difficult rather impossible to remember information concerning reactions, properties and atomic masses of elements. So we clearly need some way to organize our knowledge 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 instance if you know physical and chemical properties of one element in a group, you can predict about the physical and chemical properties of any other element present in the same group. You can use periodic table to relate trends in the reactivity of elements with their atomic structure. You can also predict which elements can form ionic or covalent bonds.

Reading

3.1 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 lawthat states **if the elements are arranged in the order of their increasing atomic numbers, their properties are repeated in a periodic manner.**

3.1.1 Periods and Groups of Elemetns.

The most commonly used form of the periodic table is shown in figure 3.1. 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.

Activity 3.1

Look at the periodic table and write number of elements present in the relevant period in the table 3.1

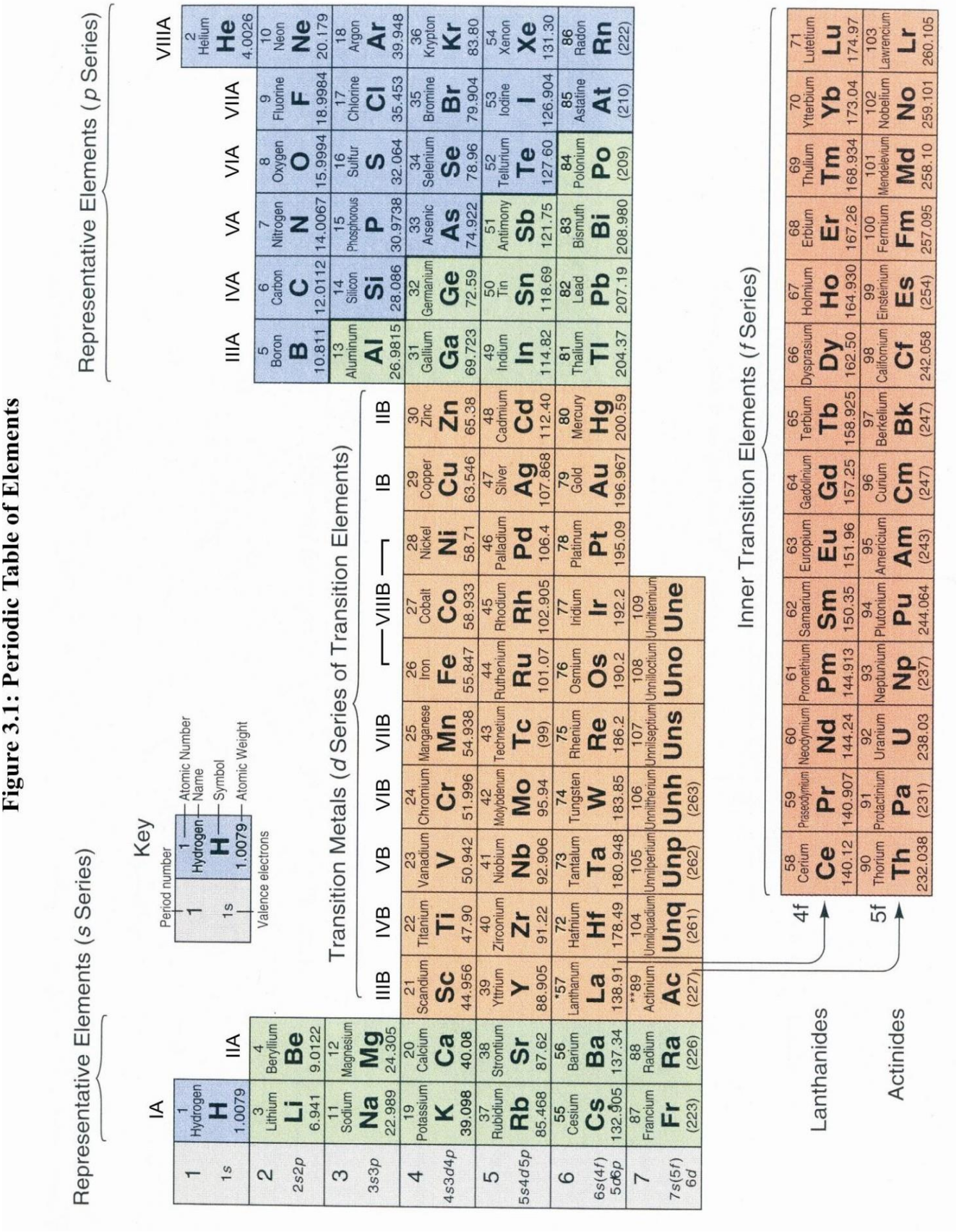
Table 3.1 Number of elements in the periods of the periodic table

Period No.	No. of elements
First	
Second	
Third	
Fourth	
Fifth	
Sixth	
Seventh	

Teacher's Point

Teacher may ask students to memorise elements of 1st two groups and 1st two periods.

Figure 3.1: Periodic Table of Elements



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.**

Elements with similar valance shell electronic configuration are placed in the same group. Each group is identified by a number and the letter A or B. Group A elements are called **normal** or **representative elements.** They are also called main group elements. Group B elements are called **transition elements.**

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

Society, Technology, Science

In 1864, John Newland, an English chemist arranged 24 elements in order of increasing atomic masses. He noticed that every eighth element, starting from any point, has similar properties. Few rows of his arrangement are shown below:

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe

His scheme however, failed because many elements were found out of place in his arrangement. For instance Ti does not resemble C and Si, Mn does not resemble N and P and Fe does not resemble O and S. However his arrangement of elements in order of increasing atomic masses formed basis for later classification of elements.

In 1869, Mendeleev, a Russian chemist developed a classification scheme of elements. He recognized that if elements were placed in order of increasing atomic masses, the properties of elements repeated at regular intervals. He arranged 65 elements in periods and groups. Development of the periodic table nicely explains how scientist can build on one another's work.

Self-Assessment Exercise 3.1

In which period and group the following elements are present in the periodic table. (a) Mg (b) Ne (c) Si (d) B

Example 3.1: Identifying the group and period of an element

Identify the group and period of $^{27}_{13}\mathrm{Al}$, $^{9}_{5}\mathrm{B}$, $^{24}_{12}\mathrm{Mg}$ on the basis of electronic configuration.

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.

Solution:

a) $^{27}_{13}\mathrm{Al} = \underset{K}{\underline{1s^2}}\ \underset{L}{\underline{2s^2,2p^6}}\ \underset{M}{\underline{3s^2,3p^1}}$

K	L	M
1	2	3

Valence shells is M

As n = 3, Al is present in the 3rd period. Since total number of electrons in the valence sub-shells are 2+1=3, it must be present in Group IIIA.

b) $^{9}_{5}B = \underset{K}{\overset{1s^2}{\underline{}}}\ \underset{L}{\overset{2s^2,2p^1}{\underline{}}}$

K	L
1	2

Valence shell is L

So n = 2, B is present in the 2nd period. Since total number of electrons in the valence shell are 2+1=3, it must be present in Group IIIA.

c) $^{24}_{12}Mg = \underset{K}{\overset{1s^2}{\underline{}}}\ \underset{L}{\overset{2s^2,2p^6}{\underline{}}}\ \underset{M}{\overset{3s^2}{\underline{}}}$

K	L	M
1	2	3

Valence shell is M

So n = 3, Mg is present in the 3rd period. Since total number of electrons in the valence shell are 2, it must be present in Group IIA.

Self-Assessment Exercise 3.2

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

(a) $^{28}_{14}Si$, (b) $^{32}_{16}S$, (c) $^{19}_{9}F$, (d) $^{40}_{18}Ar$

Example 3.2: 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:

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

Solution:

		IIA	VA	VIIA
Period 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** $3s^2 3p^5$

Self-Assessment Exercise 3.3

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

P = $1s^2 2s^2 2p^2$ Q = $1s^2 2s^2 3p^1$

R = $1s^2$ S = $1s^2 2s^2$

T = $1s^2 2s^1$ W = $1s^2 2s^2 2p^6$

X = $1s^2 2s^2 2p^6 3s^2 3p^2$ Y = $1s^2 2s^2 2p^6 3s^2 3p^6$

Z = $1s^2 2s^2 2p^1$

IA							VIIIA
	IIA		IIIA	IVA	VA	VIA	VIIA	

3.1.5 s and p Blocks in the Periodic Table

On the basis of outer most valence sub shell, elements in the periodic table can slso be classified into four blocks. Elements of Group IA and Group IIA contain their valence electrons in s sub-shell. Therefore, these elements are called **s-block** elements. The elements of Group IIIA to VIIIA (except He) are known as **p – block** elements, because their valence electrons lie in p sub-shell. Figure 3.2 shows blocks in the periodic table

3.1.6 Valence Shell Electronic Configuration and Position of an Element in the Periodic Table

You can determine the valence shell electronic configuration of an element from its position in the periodic table. Period number of element indicates n value of the valence shell. Whereas group number of element indicates the number of electrons in the valence shell.

Example 3.3: Obtaining the valence shell configuration

Write the valence shell electronic configuration of the following elements from their position in the periodic table.

(a) Phosphorus (b) Neon

Problem Solving Strategy:

Remember that

Period number = n value of valence shell

Group number = number of valence electrons

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

Solution:

a) Period number of phosphorus is 3,

As n = 3 therefore, valence shell is M

So valence electrons will be present in 3s and 3p sub-shells

The group number is 5, so there are 5 electrons in the valence shell

Two electrons will fill 3s sub-shell and remaining 3p sub-shell. Thus, the valence shell electronic configuration is $3s^2\ 3p^3$

b) Period number of Ne is 2. So, n = 2 and valence shell is L. Valence electrons will be present in 2s and 2p sub-shells.

Group number for Ne is 8,

This means there are 8 electrons in the valence shell. Two electrons will fill 2s sub-shell and remaining six 2p sub-shell. Thus the valence shell electronic configuration for Ne is $2s^2 2p^6$.

Example 3.4: Obtaining the position of element in the periodic table from lectronic configuration

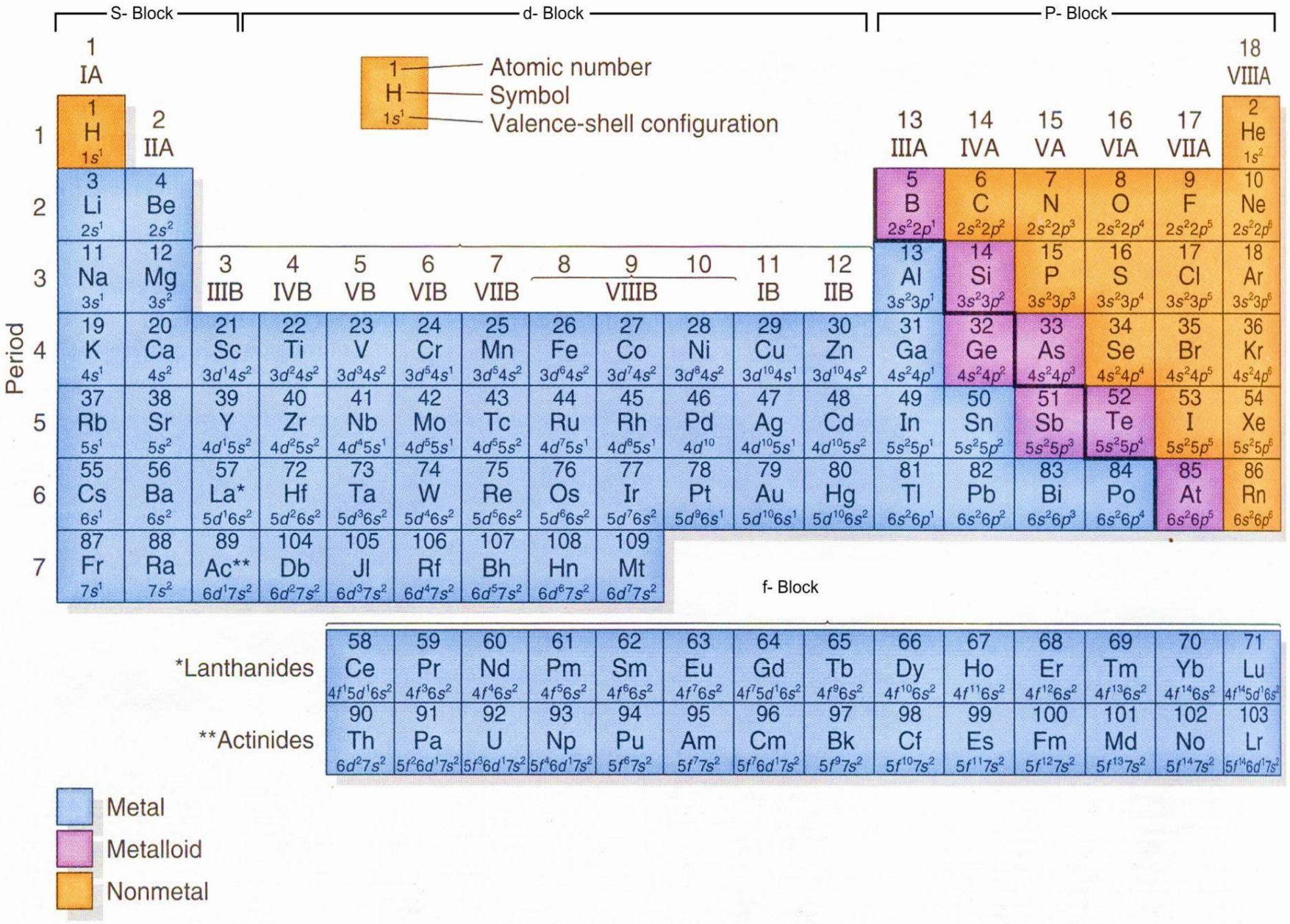
Find out the position of the following elements in the periodic table from the electronic configuration:

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

Teacher's Point

A teacher may ask student to draw Valence shells electronic configuration of some commone elements .

Figure 3.2: Blocks in the periodic table



Problem Solving Strategy:

Write electronic configuration of the element. Identify the valence shell configuration, co-efficient 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

Group number = 2 + 3=5

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

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

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

So, Period number = 2

Group number = 2 + 4 = 6

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

Self-Assessment Exercise 3.4

1. Obtain the valence shell configuration of Al and S from their position in the periodic table.
2. Find out the position of Ne (At. No 10) and Cl (At. No. 17) in the periodic table.

3.1.7 Shape of the Periodic Table

Recall that the horizontal rows in the period table are called periods. How are these periods formed?

Elements are arranged in order of increasing atomic number. First period contains only two elements, H and He. Both these elements have valence electron in K shell. K shell can not have more than two electrons. As K shell is completed at He, so the period also ends at He, Lithium (Li) atomic number 3 has one electron in L shell, so second period begins with Li. Since L shell can accommodate 8 electrons, so eight elements come in the 2^{nd} period. Second period ends at Ne which has eight electron, in L shell ($2s^2\,2p^6$).

Next elements Na has valence electron in the third shell (M – shell), in Na valence electron is present, in 3s sub – shell, which has similar electronic configuration as Li ($2s^1$), So it comes under Li. Mg with $3s^2$ valence shell electronic configuration come under Be ($2s^2$), Similarly next six elements Al, Si, P, S Cl and Ar on the bases of similar valence shell electronic configuration come under B, C, N, O, F, and Ne respectively. Ar has $3s^2\ 3p^6$ valence shell configuration similar to Ne ($2s^2\ 2p^6$). Next element K has $4s^1$ electronic configuration in the valence shell, which is similar to Na($3s^1$). So K comes under Na and a new period (4^{th}) begins with K. In this way elements having similar valance shell configuration come in the same group. The arrangement of the elements into periods has an important consequence. The elements that have similar properties end up in the same group in the periodic table.

3.2 PERIODICITY OF PROPERTIES

In section 3.1.4 you learned that, the electronic configuration of elements show a periodic variation with the increasing atomic number. Therefore, the elements also show periodic variation in their physical and chemical properties. Elements having similar valence shell electronic configuration have been placed in the same group, one below the other. Chemical properties depend on the valence shell electronic configuration. Because all the elements of a particular group have similar valence shell electronic configuration, they possess similar chemical characteristics. Physical properties depend on the sizes of atoms. Since sizes of atoms change gradually from top to bottom in a group. Therefore, elements show gradation in physical properties in the same group. In a period of periodic table the number of electrons present in the valance shell increase gradually from left to right. Their chemical and physical properties also show variation in the same manner. In this section you will learn variation in some of the physical properties of elements in a group and across a period.

3.2.1 Shielding Effect

Figure 3.2 shows electronic configuration of Li, Be and Mg.

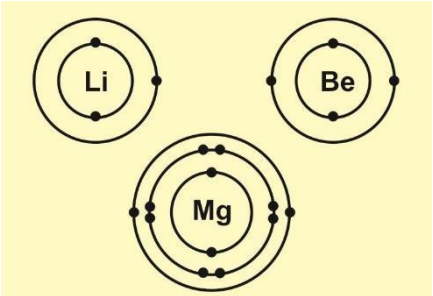


Fig 3.2 Electronic structure of Li, Be and Mg

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.

Example 3.5: 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

Teacher's Point

A teacher may ask student to draw shell of first four elements group A to show shielding effect.

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.

Self-Assessment Exercise 3.5

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 the elements depend on the electronic configurations of their atoms. Now we will discuss four properties of atoms that are influenced by the electronic configuration: atomic size, ionization energy, electron affinity and electronegativity. These properties are periodic. They generally increase and decrease in a recurring or repeating manner through the periodic table. This means they show consistent changes or trends, within a group or a period. These trends correlate with the behaviour.

3.2.2 Atomic Size

The size of an atom depends on its electronic configuration. **The size of an atom is the average distance between the nucleus of an atom and the outer electronic shell**. Figure 3.3 shows atomic radii of the main group elements.

Figure 3.3 shows the variation in atomic radii in a period and within a group. You can see two general trends in atomic radii.

(1) The atomic radius decreases in any given period as you move across the period. This is because as you move from one element to the next on its right in a period. Another electron is added to the same valence shell. At the same time positive charge on the nucleus also increases by 1. The attractive force of the nucleus for the valence shell electron increases. Therefore, the shell size and atomic radius decreases. For example, in going from lithium to beryllium, atomic size decreases. This you can understand from the valence shell electronic configuration of Li ($2s^1$) and B ($2s^2$). In going from Li to Be, there is no change in the shell number n, but atomic number increases from 3 to 4. Due to this the force of the nucleus for the valence shell electron increases. Therefore, atomic radius decreases.

The atomic radius increases in any given main group as you move down the group of elements. This is because the size of an atom is determined by the size of its valence shell. As you move to the next lower element in the group, the atom has an additional shell of electrons. This increases atomic radius. For example, in going from Li to Na atomic radius increases.

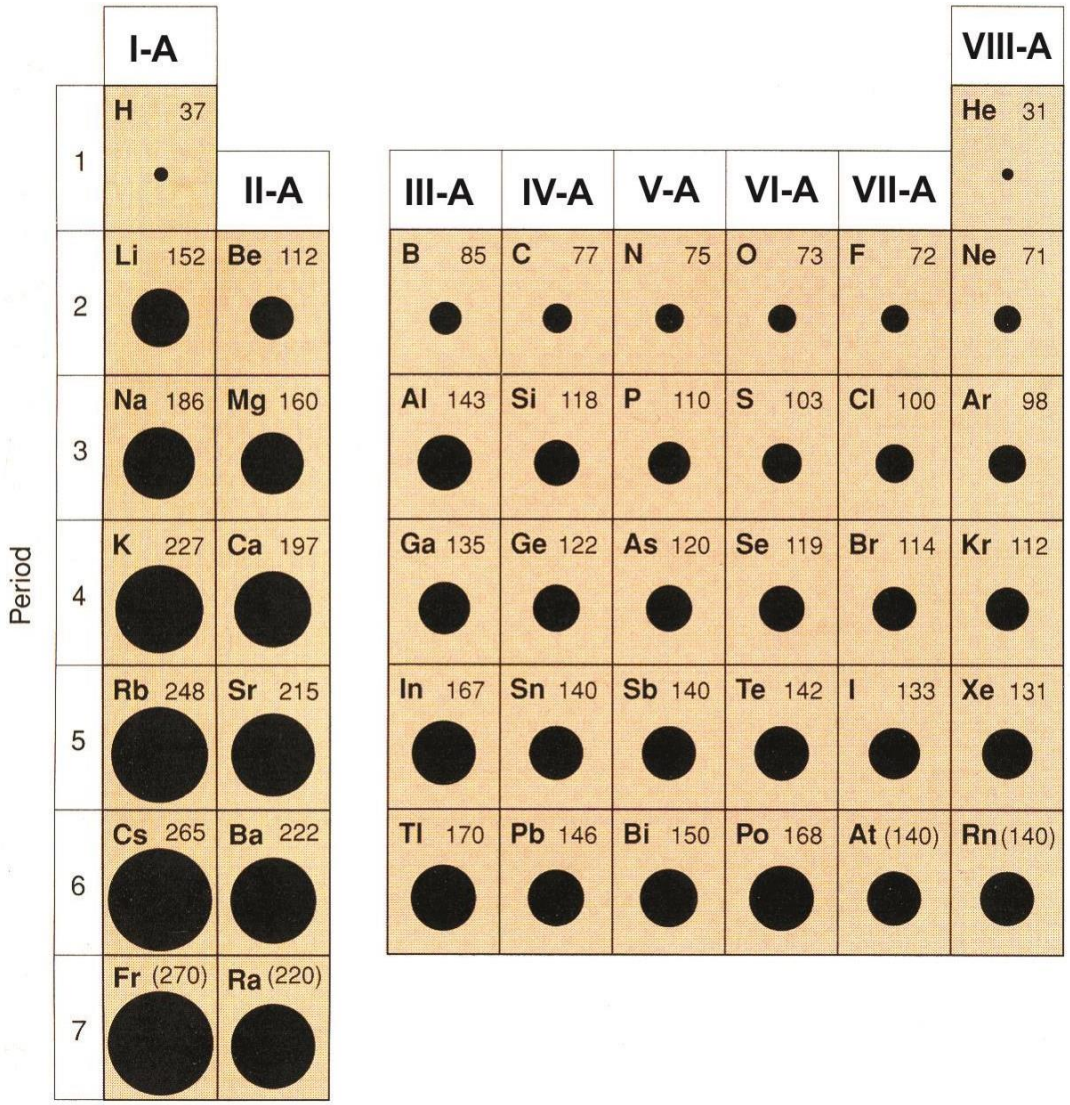


Figure 3.3: Atomic radii of the main group elements (in picometer)

Consider electronic configuration of Li ($1s^2\ 2s^1$) and Na($1s^2$, $2s^2$, $2p^6$, $3s^1$). A new electronic shell has been added that increases atomic size.

Example 3.6: 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

Self-Assessment Exercise 3.6

Using the periodic table but without looking at the figure 3.3, 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

3.2.3 Ionization Energy

Period	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 3.4 Ionization energies of the main group elements

You have learned in section 1.3.1 how cations are formed. 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"**.

$$M(g) + \text{ionization energy} \downarrow\downarrow M^{+}(g) + e^{-}$$

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 3.4 shows the ionization energies of the main group elements. Values are given in units of kJ/mole^{-1} or kJ/mole.

Trends in the values of ionization energies.

The ionization energy value decreases from top to bottom in a group. This is because the shielding effect in atoms increases as you descend. Greater shielding effects results in a weaker attraction of the nucleus for the valence electrons. So, they are easier to remove. This leads to decrease in ionization energy from top to bottom in a group.

Which atom has greater shielding effect, Li or Na?

As you move from left to right in a period, the shielding effect remains constant. But progressively nuclear charge increases. A stronger force of attraction between nucleus and the

valence electron increases. This leads to increase in ionization energy from left to right in a period.

Which atom has higher ionization energy, Li or Be?

Example 3.7: 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.

Self-Assessment Exercise 3.7

Which atom has the smaller ionization energy?

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

3.2.4 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 uninegative gaseous ion**.

$$X(g) + e^- \downarrow\downarrow X^-(g) + \text{electron affinity}$$

Figure 3.5 shows electron affinities of main group elements.

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

The electron affinity decreases from top to bottom in a group. This is due to increase in shielding effect. Due to increase in shielding effect 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 3.5 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 grade XI.

3.2.5 Electronegativity

Electronegativity is the ability of an atom to attract the electrons towards itself in a chemical bond. Figure 3.6 shows as scale of electronegativities of the elements devised by Linus Pauling. The American chemist Linus Pauling devised a method for calculating 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 3.6 the electronegativities of elements.

Activity 3.3

Determining the general trends in the electronegativities

You will need:

- Figure 3.6

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 electronegativites across a period and write reason.
4. Move from top to bottom in Groups IA and IIA and note down the variation in electronegativites value.
5. Move from top to bottom in Groups VIA and VIIA and note down the variation in electronegativities value.
6. Make generalization about the trend in electronegativity values in a group. Give reason.

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 group IIIA to VIIIA 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 uninegative gaseous ion.

REFERENCES FOR ADDITIONAL INFORMATION

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

Review Questions

1. Encircle the correct answer:

(i) Number of periods in the periodic table are:
(a) 8 (b) 7 (c) 16 (d) 5

(ii) Which of the following groups contain alkaline earth metals?
(a) 1A (b) IIA (c) VIIA (d) VIIIA

(iii) Which of the following elements belongs to VIIIA?
(a) Na (b) Mg (c) Br (d) Xe

(iv) Main group elements are arranged in_____groups.
(a) 6 (b) 7 (c) 8 (d) 10

(v) Period number of $^{27}_{13}Al$ is:
(a) 1 (b) 2 (c) 3 (d) 4

(vi) Valence shell electronic configuration of an element M (atomic no. 14) is:
(a) $2s^2 2p^1$ (b) $2s^2 2p^2$ (c) $2s^2 2p^3$ (d) $3s^2 3p^2$

(vii) Which of the following elements you expect to have greater shielding effect?
(a) Li (b) Na (c) K (d) Rb

(viii) As you move from right to left across a period, which of the following does not increase:
(a) electron affinity (b) ionization energy
(c) nuclear charge (c) shielding effect

(ix) All the elements of Group IIA are less reactive than alkali metals. This is because these elements have:
(a) high ionization energies (b) relatively greater atomic sizes
(c) similar electronic configuration (d) decreased nuclear charge

2. Give short answers

i) Write the valence shell electronic configuration of an element present in the 3rd period and Group IIIA.
ii) Write two ways in which isotops of an element differ.
iii) Which atom has higher shielding effect, Li or Na?
iv) Explain why, Na has higher ionization energy than K?
v) 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.**

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

9. **Write the valence shell electronic configuration for the following groups:**

a. Alkali metals
b. Alkaline earth metals
c. Halogens
d. Noble gases

10. **Write electron dot symbols for an atom of the following elements**

(a) Be (b) K (c) N (d) I

11. **Write the valence shell electronic configuration of the atoms of the following elements.**

(a) An element present in period 3 of Group VA

(b) An element present in period 2 of Group VIA

12. Copy and complete the following table:

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

13 Imagine you are standing on the top of Neon-20 nucleus. How many kinds of sub-atomic particles you would see looking down into the nucleus and those you would see looking out from the nucleus.

14. Chlorine is a reactive element used to disinfect swimming pools. It is made up of two isotopes Cl-35 and Cl-37. Because Cl-35 is more than Cl-37, the atomic mass of chlorine is 35.5amu. is closer to 35 than 37. Write electronic configuration of each isotope of chlorine. Also write symbol for these isotopes (atomic number for chlorine is 17).

15. In which block, group and period in the periodic table where 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) $1s^2$

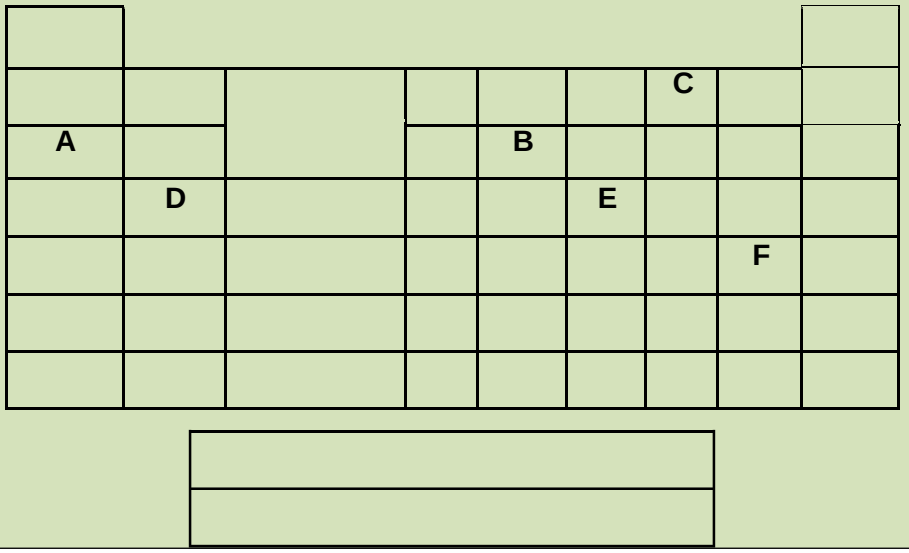
Think-Tank

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

2. 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 2080kJ/mole and that of the other is 496kJ/mole. Match each ionization energy with one of the given electronic configuration. Give reason for your choice.

3. 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.

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



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 dipositve 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?

5. 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

6. Argue in what region of the periodic table you will find elements with relatively

a) high ionization energies
b) low ionization energies