UNIT 2

Periodic Classification of the Elements

Unit Outcomes

After completing this unit, you will be able to:

- *understand the periodic classification of the elements;*
- *develop skills in correlating the electron configuration of elements with the periodicity of the elements, and in predicting the trends of periodic properties of elements in the periodic table;*
- *appreciate the importance of classification in chemistry; and*
- *demonstrate scientific inquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, measuring, asking questions, interpreting illustrations, drawing conclusions, applying concepts, and problem solving.*

MAIN CONTENTS

Start-up Activity

How can you use a table of repeating events to predict the next events? Table 2.1 shows a familiar table of repeating properties. What is it?

Of course, it is a calendar, but it is a calendar with a difference – it is missing some information. You can determine what is missing and fill in the blanks.

1. Look at the calendar again. The calendar has columns of days, Sunday through Saturday. The calendar also has horizontal rows. They are its weeks.
2. Examine the information surrounding each empty spot. Can you tell what information is needed in each empty spot?
3. Fill in the missing information. For example:

Conclude and Apply

1. One day in column 3 is marked X, and a day in column 4 is marked Y. What dates belong to these positions? Discuss your answers in your group.
2. Column 5 does not have a name. What is the correct name of this column?
3. What dates are included in the third row of the table?
4. Assuming that the previous month had 30 days, what day would the 28th of that month have been? What row of this table would it appear in?
5. How do you relate this periodicity with the periodic classification of elements?

Table 2.1 Periodicity in time.

Sun	Mon	Tues	Wed		Fri	Sat
				1	2	3
4	5	6	7	8	9	10
11	12	X	Y	15	16	17
18	19	20	21	22	23	24
25	26	27	28	29	30	31

2.1 INTRODUCTION

Competencies

By the end of this section, you will be able to:

- Describe periodicity.

Activity 2.1

Form a group and discuss the following concepts. Present your discussion to the rest of the class.

What was the basis for the early attempts in classifying the elements? Outline the contributions of some scientists you have read about in classfying the elements.

Before the beginning of the 18^{th} century, it was easy to study and remember the properties of the elements because very few were known. However, in the middle of the 19^{th} century, many more elements were discovered. Scientists then began to investigate possibilities for classifying the known elements in a simple and useful manner. After numerous attempts, the scientists were ultimately successful. They grouped elements with similar properties together. This arrangement is known as the classification of elements.

Early Attempts in classifying the elements

What is meant by the term periodicity? Can you give some periodic events in nature?

Early attempts to classify elements were based merely on atomic mass. Then scientists began to seek relationships between atomic mass and other properties of the elements.

i) **Dobereiner's Triads:** One of the first attempts to group similar elements was made in 1817 by the German chemist J. Dobereiner. He put together similar elements in group of three or triads. According to Dobereiner, when elements in a triad are arranged in the order of increasing atomic masses, the middle element had the average atomic mass of the other two elements. For example, because the atomic mass of bromine is nearly equal to the average atomic mass of chlorine and iodine, he considered these three elements to constitute a triad when arranged in this order: chlorine, bromine, iodine see Table 2.2.

Table 2.2 Dobereiner's Triads.

Triads	Atomic Masses
Chlorine	35.5
Bromine	80
Iodine	127
Average atomic masses of chlorine and iodine $= \dfrac{35.5+127}{2} = 81.25$	

Reading Check

Do you think that this classification works for all elements?

ii) **Newlands's Law of Octaves:** In 1864, John Newlands, an English chemist, reported the law of octave, which is also known as the law of eight. He stated that when elements are arranged in increasing order of their atomic masses, every eighth element had similar properties to the first element.

Newlands first two octaves of eight elements are shown below:

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

However, the law of octaves could not be applied beyond calcium.

Reading Check

With the aid of an encyclopedia, reference books or other resources, write a report on:

a J. Dobereiner and

b J. Newlands works in organizing the elements.

In your report include the merits and demerits of their works.

Exercise 2.1

1. Decide whether the principle of Dobereiner's triad can be applied in the following groups of three elements.

   a Be, Ca, Sr   b Li, Na, K

2. Newlands stated that there was a periodic similarity in properties of every eighth element in his system. However, today we see that for periods 2 and 3, the similarity occurs in every ninth element. What is the reason? Explain.

2.2 The Modern Periodic Table

Competencies

By the end of this section, you will be able to:

- state Mendeleev's and modern periodic law;
- describe period and group;
- explain the relationship between the electronic configuration and the structure of the modern periodic table;
- describe the three classes of the elements in the modern periodic table;
- explain the four blocks of the elements as related to their electronic configuration in the modern periodic table;

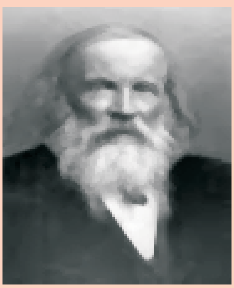
- tell the block of an element from its electronic configuration;
- give group names for the main group elements;
- classify the periods into short, long, and incomplete periods;
- tell the number of groups and periods in the modern periodic table;
- tell the number of elements in each period;
- predict the period and group of an element from its atomic number; and
- tell the block and group of an element from its electronic configuration.

Activity 2.2

Form a group and perform the following tasks. Share your findings with the rest of the class.

1. Make a list of the symbols for the first eighteen elements. Beside each symbol, write its electronic configuration.
2. Draw sets of vertical boxes. In order of increasing atomic number, fill into each set the symbol of all elements having the same outer electron configurations. How many sets are there? Record your answer.
3. Draw sets of horizontal boxes. In order of increasing atomic number, fill into each set the symbols of all elements having the same number of shells. How many sets are there? How many elements are there in each set? Record your answers.
4. Do you see any regular patterns that you created in Steps 2 and 3?
5. Draw one complete table which shows all elements;
   a with the same number of outermost electrons in a vertical column
   b filling the same outer electron shell in a horizontal row.

Historical Note



Dimitri Mendeleev

While attempting to group the elements according to their chemical properties and atomic weights, Dimitri Mendeleev developed the periodic table and formulated the periodic law. Because his classification revealed recurring patterns (periods) in the elements, Mendeleev was able to leave spaces in his table for elements that he correctly predicted would be discovered.

2.2.1 The Periodic Law

A. Mendeleev's Periodic Law

In 1869, the Russian chemist Dimitri Mendeleev and the German scientist Luthar Meyer independently published periodic arrangements of the elements based on increasing atomic mass.

Mendeleev observed that, when elements are arranged according to increasing atomic mass, the chemical and physical properties of the elements recur at regular intervals. This periodic variation and the recurrence of the properties of the elements led to the formulation of the Periodic Law.

Mendeleev's periodic law states that the properties of the elements are periodic functions of their atomic masses. Only 63 elements were known when Mendeleev constructed his table in 1871. Mendeleev organized his table in columns, with each column containing elements that have similar chemical properties. Accordingly, elements in the same column gave family or groups of elements.

How did Mendeleev know where to leave gaps for undiscovered elements in his periodic table?

Mendeleev left blank spaces for the undiscovered elements and also predicted masses and other properties of these unknown elements almost correctly. For example, neither gallium nor germanium were discovered when Mendeleev constructed his periodic table. But he predicted the existence and properties of these unknown elements.

Mendeleev left two blank spaces for these two elements in the table, just under aluminium and silicon. He called these unknown elements 'eka-aluminium' and 'eka-silicon.' ('eka' means 'first'). What he meant by 'eka-aluminium' is "a currently known element (galium) following aluminium".

Later on, in 1874, the element Gallium (eka-aluminium in Mendeleev's system) was discovered. In 1886, the element Germanium (eka-silicon) was discovered. The observed properties of these elements were remarkably very close to those in Mendeleev's predictions. Table 2.3 shows the properties of eka-silicon predicted by Mendeleev and compares them to the observed properties of Germanium.

Table 2.3 Comparison of Mendeleev's predictions for the properties of Eka-silicon with Germanium.

Property	Mendeleev's Predictions for eka-silicon (Es) in 1871	Observed Properties for Germanium (Ge) in 1886
Atomic mass	72	72.6
Density (g/cm^3)	5.5	5.47
Colour	Dark Gray	Light Gray
Oxide formula	EsO_2	GeO_2
Density of oxide (g/cm^3)	4.7	4.7
Chloride formula	$EsCl_4$	$GeCl_4$
Density of chloride (g/cm^3) . . .	1.9	1.887
Boiling point of chloride	< 100°C	86°C

Defects in Mendeleev's periodic table

1. Position of isotopes: The isotopes were not given separate places in Mendeleev's periodic table. Since elements are arranged in order of increasing atomic masses, the isotopes belong to different groups (because *isotopes have different masses*).
2. Wrong order of atomic masses of some elements: When certain elements are grouped on the basis of their chemical properties, some elements with higher atomic masses precede those with lower atomic masses. For example, argon, with atomic mass of 39.95, precedes potassium with atomic mass of 39.1.

B The Modern Periodic Law

What was Mosley's contribution to the modern form of the Periodic Table?

In 1913, the English physicist Henry Mosley determined the atomic number of each of the elements by analyzing their X-ray spectra. He observed that when each element was used as a target in an X-ray tube, it gave out X-rays with a characteristic wavelength. The wavelength depends on the number of protons in the nucleus of the atom and was constant for a given element. By arranging the elements in order of decreasing wavelength, Mosley was able to assign atomic number to each element. The atomic number of every element is fixed, and it clearly distinguishes one element from another. '*No two elements can have the same atomic number*.' For example, atomic number 8 identifies the element oxygen. No other element can have atomic number 8.

Therefore, the atomic number of an element is the fundamental property that determines the chemical behavior of the element. The discovery of atomic number led to the development of the modern periodic law. The modern periodic law states that: "the properties of the elements are periodic function of their atomic numbers." This means that when elements are arranged according to increasing atomic number, elements with similar physical and chemical properties fall in the same group.

2.2.2 Characteristics of Groups and Periods

Many different forms of the periodic table have been published since Mendeleev's time. Today, the long form of the periodic table, which is called the modern periodic table, is commonly in use. It is based on the modern periodic law. In the modern periodic table, elements are arranged in periods and groups.

What are the basis for classifying the elements into groups and periods?

***What are the similarities and differences in the electron configuration of* S *and* Cl?**

Periods: The horizontal rows of elements in the periodic table are called periods or series.

Elements in a period are arranged in increasing order of their atomic numbers from left to right.

There are 7 periods in the modern periodic table, and each period is represented by an Arabic numeral: 1, 2 . . . and 7.

- Elements in the same period have the same number of shells.
- Periods 1, 2, and 3 are called short periods while periods 4, 5, and 6 are known as long periods.
- Period 1 contains only 2 elements, hydrogen and helium. Period 2 and period 3 contain 8 elements each.
- Period 4 and period 5 contain 18 elements each. Period 6, the longest period, has 32 elements. Period 7, which is an incomplete period, contains more than 24 elements.Period 7 element is radioactive and/or an artificial element.
- Except for the first period, all periods start with an alkali metal and ends with a noble gas.

Table 2.4 The number of elements in a given period and the orbitals being filled.

Period number	Orbitals occupied	Number of elements
1	1*s*	2
2	2*s*, 2*p*	8
3	3*s*, 3*p*	8
4	4*s*, 3*d*, 4*p*	18
		and so on

The position of an element in a given period can be determined by the number of shells occupied with its electrons. Accordingly, the number of shell is equal to the number of period to which the element belongs.

Example

Electronic configuration of $^{23}_{11}Na$ = $1s^2 2s^2 2p^6 3s^1$ (2, 8, 1). Sodium has 3 main shells. Hence, sodium is found in period 3.

***What are the similarities and differences in the electron configuration of* F *and* Cl?**

Groups or families: are the vertical columns of elements in the periodic table. There are 18 columns or groups in the modern periodic table.

- Group numbers are usually designated with the Roman numerals I to VIII each followed by the letter A or B.

  These are: IA . . . . VIII A Main groups (A groups)

  IB . . . . VIIIB Sub groups (B groups)

- Elements in a given group have the same number of outermost shell electrons.
- Elements in the same group have similar chemical properties.
- For the main group elements the group number equals the number of valence electrons.

Example

Electronic configuration of $^{35}_{17}Cl$ = $1s^2 2s^2 2p^6 3s^2 3p^5$ (2, 8, 7). The number of valence electrons of chlorine is 7. Hence chlorine is found in Group VIIA.

PERIODIC TABLE

Relative Atomic Mass → 1.0079
H ← Symbol
Atomic Number → 1

Legend

- Alkali Metals
- Alkaline Earth Metals
- Transition Metals
- Rare Earth metals
- Nonmetals
- Noble gases
- Other Metals
- Metalloids

Group → ; Period ↓

Period	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1.0079 H 1																	4.003 He 2
2	6.941 Li 3	9.01 Be 4											10.81 B 5	12.01 C 6	14.01 N 7	15.999 O 8	18.998 F 9	20.18 Ne 10
3	22.990 Na 11	24.31 Mg 12											26.98 Al 13	28.09 Si 14	30.97 P 15	32.06 S 16	35.45 Cl 17	39.95 Ar 18
4	39.098 K 19	40.08 Ca 20	44.96 Sc 21	47.90 Ti 22	50.94 V 23	51.996 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.70 Ni 28	63.55 Cu 29	65.37 Zn 30	69.72 Ga 31	72.59 Ge 32	74.92 As 33	78.96 Se 34	79.90 Br 35	83.80 Kr 36
5	85.458 Rb 37	87.62 Sr 38	88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	(98) Tc 43	(101.07) Ru 44	102.91 Rh 45	106.40 Pd 46	107.87 Ag 47	112.47 Cd 48	114.82 In 49	118.69 Sn 50	121.75 Sb 51	127.60 Te 52	126.90 I 53	131.30 Xe 54
6	132.91 Cs 55	137.33 Ba 56	136.91 La★ 57	178.49 Hf 72	180.95 Ta 73	183.85 W 74	186.21 Re 75	190.20 Os 76	192.22 Ir 77	195.09 Pt 78	196.97 Au 79	200.59 Hg 80	204.37 Tl 81	207.19 Pb 82	208.98 Bi 83	(209) Po 84	(210) At 85	(222) Rn 86
7	(223) Fr 87	226.03 Ra 88	227.03 Ac* 89	(261) Rf 104	(262) Ha 105	(266) Sg 106	(262) Bh 107	(277) Hs 108	(268) Mt 109	(281) Uun 110	(272) Uun 111	(285) Uub 112						

★ Lanthanides	40.12 Ce 58	140.91 Pr 59	144.24 Nd 60	(145) Pm 61	(145) Sm 62	151.96 Eu 63	157.25 Gd 64	158.93 Tb 65	162.50 Dy 66	164.93 Ho 67	167.26 Er 68	168.93 Tm 69	168.93 Yb 70	174.97 Lu 71
* Actinides	232.04 Th 90	231.04 Pa 91	238.03 U 92	237.05 Np 93	(244) Pu 94	(243) Am 95	(247) Cm 96	(247) Bk 97	(251) Cf 98	(252) Es 99	(257) Fm 100	(260) Md 101	(259) No 102	(262) Lr 103

Exercise 2.2

Give appropriate answers for the following questions.

1. Which group of elements was missing from Mendeleev's periodic table as compared to the modern periodic table?
2. Why does the first period contain only two elements?
3. To which group and period do the following elements belong?

   a carbon
   b neon
   c aluminium
   d potassium
   e calcium
   f sulphur

2.2.3 CLASSIFICATION OF THE ELEMENTS

Activity 2.3

Perform the following tasks in groups and present your conclusion to the class.

For the following elements, determine the valence electrons, and identify the sub–shell (*s*, *p*, *d* or *f*) in which the last electron of each element enters.

- a Nitrogen (atomic number = 7)
- b Sodium (atomic number = 11)
- c Silicon (atomic number = 14)
- d Iron (atomic number = 26)
- e Zinc (atomic number = 30)
- f Krypton (atomic number = 36)
- g Cerium (atomic number = 58)

Elements in a periodic table can be classified into three distinct categories based on their electron configuration and the type of sub-level being filled. These are the representative elements, the transition elements, and the rare-earth elements.

1. Representative Elements: *s*-and *p*-block elements

These are elements in which valence electrons are filling the *s*- or *p*-orbitals. Representative elements are also known as Main Group Elements. They include elements in groups IA through VIIIA.

s-block elements are elements in which the last electron enters the *s*-orbital of the outermost shell. Their general valence shell configuration is ns^1 and ns^2, where n represents the outermost shell. They are found on the left side of the periodic table and contain the first two groups: Group IA and Group IIA. These two groups contain very reactive metals.

p-block Elements are elements in which the last electron enters the *p*-orbital of the outermost shell. They have the general valence shell electron configuration of ns^2np^1–ns^2np^6. *p*-block elements are located at the right hand side of the periodic table and they include six groups-Group IIIA to VIIIA. Most of these elements are non-metals.

Table 2.5 Valence electron configuration, blocks and common names for the representative elements.

Group Number	Common name	General valence shell electron configuration		Number of valence electrons
IA	Alkali metals	ns^1	*s*-block	1
IIA	Alkaline earth metals	ns^2		2
IIIA	Boron family	ns^2np^1	*p*-block	3
IVA	Carbon family	ns^2np^2		4
VA	Nitrogen family	ns^2np^3		5
VIA	Oxygen family (Chalcogen)	ns^2np^4		6
VIIA	Halogens	ns^2np^5		7
VIIIA	Noble gases (Inert gases)	ns^2np^6	except He	8

Why is helium classified as a noble gas?

2. Transition Elements: *d*-block elements

These elements are found in the periodic table between the s-block and p-block elements. In these elements, the valence electrons are being added to the *d*-orbital of the outermost shell. They are also known as *d*-block elements. They are designated as 'B' group, and they consist of groups IB – VIIIB. The transition elements are found in periods 4, 5, 6, and 7. They include common elements such as iron, gold, and copper. *d*-block elements are also called transition metals.

3. Rare-Earth Elements: *f*-block elements

These elements are also known as inner-transition elements. They are elements in which the last electrons are being added to the *f*-orbital of the outermost shell. The inner transition elements are in periods 6 and 7. The period 6 inner transition series fills the 4*f*-orbitals. It is known as the Lanthanide series because it occurs after the element lanthanum and its elements are similar to it. The period 7 inner transition series fills the 5*f*-orbitals and appears after the element actinium. Its elements are similar to actinium, thus it is known as the Actinide series.

Figure 2.1 presents a scheme to show the basic structure of the periodic table. On the basis of the electron configuration of the elements the table is divided into *s*, *p*, *d* and *f*-blocks.

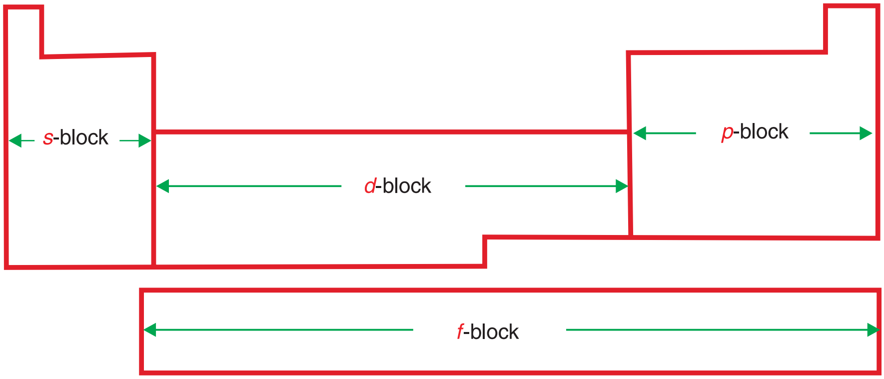


Figure 2.1 Division of the modern periodic table.

Reading Check

In the periodic table, you have seen a section called rare earth elements. Find this term in an encyclopedia, in reference books, or in other resource material. Read and analyze the information you find. Write a report on these elements.

Exercise 2.3

Give appropriate answers for the following questions:

1. For the following elements, write their electron configurations and determine the group and period number of each element.

   a Na, Ca, Al   b Cl, S, Ar

2. Bromine is a Group VIIA and period-4 element. What is the valence shell configuration of bromine?
3. Deduce the group, period, and block of the elements with atomic numbers:

   a 37   b 24   c 32

Critical Thinking

4. Why is hydrogen, a non-metal, usually placed with Group I elements in the periodic table, even though it does not show a metallic property like the alkali metals?

2.3 PERIODIC PROPERTIES IN THE PERIODIC TABLE

Competencies

By the end of this section, you will be able to:

- explain the general trends in properties of the elements as they move down a group of the periodic table;

- explain the general trends in properties of elements across a period;
- deduce the properties of an element from its position in the periodic table; and
- make a chart to show the trends in properties of elements in the periodic table.

Activity 2.4

Refer any chemistry book and look up values for atomic size, ionization energy, electronegativity and other properties of elements. Enter the values in the periodic table, and see if there any obvious trends in the properties as you go across a period or down a column of the table. Report your findings to the class.

In the periodic table the properties of the elements such as atomic size, ionization energy, electron affinity, and electronegativity show a regular variation with in a group or across a period.

2.3.1 Periodic Properties within a Group

Elements in the same group have the same number of valence electrons and also exhibit similar chemical properties. For example, group IA elements have 1 valence electron; those in Group IIA have 2 valence electrons, and so on. Generally, it is possible to conclude that the number of valence electrons determines the group number of an element.

Table 2.6 Electron configuration and number of valence electrons of Group IA elements.

Element	Electron Configuration	Number of Valence electron	Group Number
Li	2, 1	1	
Na	2, 8, 1	1	
K	2, 8, 8, 1	1	IA
Rb	2, 8, 18, 8, 1	1	
Cs	2, 8, 18, 18, 8, 1	1	

The periodic properties of the elements can be explained on the basis of nuclear charge and effective nuclear charge.

Nuclear charge (Z): is the total positive charge in the nucleus of an atom.

Effective Nuclear charge (Z_{eff}): In an atom, the outermost shell electrons (or valence electrons) are attracted to the nucleus and simultaneously repelled by the inner shell electrons. The attraction of the nucleus for the valence electrons is also reduced because inner electrons shield (or screen) the valence electrons. As a result, these inner electrons reduce the attraction of the nuclear charge. The resulting net-positive nuclear charge attracting the valence electrons is called effective nuclear charge, Z_{eff}.

Effective nuclear charge relates the nuclear charge to the number of shells (size) of the atom.

The effective nuclear charge is the difference between the nuclear charge (Z) and the inner electrons (S) that shield the valence electrons. The effective nuclear charge is always less than the actual nuclear charge.

$$Z_{eff} = Z - S$$

Let us consider the sodium atom with its 11 electrons: $1s^2 2s^2 2p^6 3s^1$ (2, 8, 1). The 10 inner electrons of sodium completely cancel the 10 units of nuclear charge on the nucleus. In this way, the inner electrons shield the valence electrons from the full attractive force of the nucleus and leave an effective nuclear charge of + 1 (Z_{eff} = + 1).

The fact that the inner electrons shield or screen the outer electrons from the full charge of the nucleus is known as shielding or screening effect.

Activity 2.5

Form a group and perform the following task:

Draw Bohr's model for the following elements and indicate the shielding shells, shielding electrons and effective nuclear charge:

a beryllium, magnesium and calcium;

b lithium, carbon and fluorine.

Present your findings to the class.

The following properties of elements vary in a regular periodic manner.

1. Nuclear Charge

On moving down a given group of the periodic table, nuclear charge progressively increases, but effective nuclear charge remains nearly constant.

2. Atomic Size/Atomic Radius

Where in a group do you find atoms with the largest atomic radius? Why?

It is difficult to measure the size of an atom directly. The electron cloud enveloping the nucleus does not have a clear boundary because its electrons do not have fixed distances from the nucleus. Therefore, atoms do not have definite outer boundaries.

The size of an atom is defined in terms of its atomic radius. For metals, atomic radius is defined as one-half the distance between the nuclei of the two adjacent atoms. For elements that exist as diatomic molecules (such as chlorine), atomic radius is equal to one-half of the distance between the nuclei of the atoms in the molecule. Figure 2.2 illustrates the atomic radius of chlorine in a chlorine molecule.

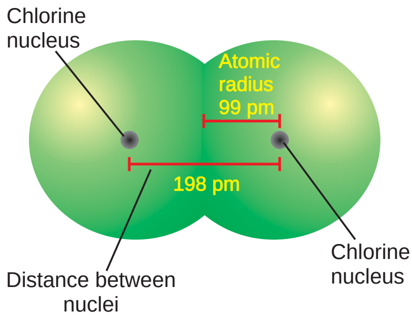


Figure 2.2 The representation of atomic radius in a chlorine molecule.

In moving down a group, atomic radius of the elements mainly depends on the number of shells.

Activity 2.6

Form a group and perform the following task; present your findings to the class.

The following values are given for atomic radii (in Å) of group IA elements:

1.54, 1.34, 2.35, 2.16 and 1.96

Based on the given information;

1. fill the following table with the number of shells and the appropriate values for atomic radii corresponding to the symbols of the elements.
2. explain the reason for the observed trend.

Elements	Number of Shells	Atomic Radius (Å)
Li		
Na		
K		
Rb		
Cs		

3. Ionization Energy

In which region of the periodic table do you find elements with the:

a lowest tendency to lose electrons, and

b highest tendency to lose electrons?

Ionization energy is the minimum energy required to remove the outermost shell electron from an isolated gaseous atom or ion.

Ionization energy is represented by the following equation (*where* M *denotes any metal*).

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

The electrons in an atom can be successively removed, one after another. Thus, the first ionization energy is the energy needed to remove the first valence electron, the second ionization energy is the energy needed to remove the second valence electron, and so on.

$M\ (g) + \text{energy} \longrightarrow M^{+}\ (g) + e^{-}$ ⇒ *First ionization energy*

$M^{+}\ (g) + \text{energy} \longrightarrow M^{2+}\ (g) + e^{-}$ ⇒ *Second ionization energy*

For a given element, the second ionization energy is higher than the first one.

Ionization energy is always a positive value (and therefore is an endothermic process) because energy is required to remove an electron from an atom. Ionization energy is measured in electron volts (eV) or kiloJoules per mole (kJ/mol).

Ionization energy is a measure of the tendency of an atom to lose an electron. Metals easily lose electrons and thus have low ionization energy. Non-metals have high ionization energy because they do not easily lose electrons.

Activity 2.7

Form a group and perform the following task. Present your findings to the class.

Rank each set of the following elements in order of decreasing ionization energy and explain the trend in ionization energy of the elements down a group.

i) Ca, Sr, Mg, Be

ii) K, Li, Rb, Na

iii) Cl, F, I, Br

Generally, ionization energy is affected by the following factors:

i) Atomic size: As atomic size increases, the valence electrons are less tightly held by the nucleus. Thus, less energy is required to remove these electrons. For example, the energy needed to remove an electron from a cesium atom is lower than from a lithium atom.

ii) Effective nuclear charge: The smaller the effective nuclear charge of an atom, the lower is the energy needed to remove an electron from the atom.

iii) Types of electrons: The closer an electron is to the nucleus, the more difficult it is to remove the electron. In a given energy level, *s*-electrons are closer to the nucleus than *p*-electrons. Similarly, *p*-electrons are closer than *d*-electrons, and *d*-electrons are closer than *f*-electrons. Hence, ionization energy decreases in the order of: $s > p > d > f$.

iv) Screening effect by the inner electrons: As described earlier, inner shell electrons shield the valence electrons from the nuclear charge. The more inner electrons there are, the higher the screening effect, and therefore the easier it is to remove the valence electrons. Screening decreases ionization energy.

v) Electron configuration (stability): It is easier to remove electrons from unstable sublevels than from stable ones. Half-filled (p^3, d^5, f^7) and completely-filled (d^{10}, p^6, f^{14}) sublevels are more stable. For example, more energy is required to remove a p^3 electron than a p^4 electron. As a result of it, the first ionization energy of nitrogen is higher than that of oxygen.

Generally, with the increasing atomic number, the first ionization energy decreases down the same group.

Activity 2.8

After studying the given information, identify the elements represented by X, M, and Y; discuss the findings in your group, and then share your conclusion with the other groups.

Element X

- has a relatively high ionization energy.
- generally forms an ion with a – 2 charge.
- has an outermost electron configuration of $3s^2 3p^4$.

Element M

- reacts with oxygen to form M_2O.
- has a very low ionization energy.
- is in the fourth period of the periodic table.

Element Y

- is a transition element.
- is used in Ethiopian coinage.
- has 10 electrons in the 3d orbital

4. Electron Affinity (E_A)

To which group it is easier to add electrons; alkali metals or halogens?

Non-metals gain electrons and therefore form negative ions. The tendency of an atom to form a negative ion is expressed in terms of electron affinity (E_A).

Electron affinity is defined as the energy released in kilojoules/mole, when an electron is added to an isolated gaseous atom to form a gaseous ion. It is a measure of the attraction or 'affinity' of the atom for the extra added electron.

$$X(g) + e^- \longrightarrow X^- (g) + \text{energy}$$

Since energy is liberated during the process, electron affinity is expressed as a negative value. For example, when an electron is added to a fluorine (F) atom, 328 kJ/mol of energy is released to produce a fluoride ion (F^-), and E_A is – 328 kJ/mol.

$$F(g) + e^- \longrightarrow F^- (g) + E_A = -\ 328 \text{ kJ/mol}$$

Electron affinity is a measure of the strength of an atom to attract an additional electron. The smaller is the atomic size of an element, the stronger is the tendency to form negative ions, and consequently the higher the electron affinity. Generally, electron affinity depends on atomic size and effective nuclear charge of the elements.

Activity 2.9

Form a group and discuss the following concepts.

1. Why does the electron affinity of Cl is higher than that of F?
2. Explain why noble gases have extremely low (almost zero) electron affinities?
3. Explain why halogens have the highest electron affinities?

Present your findings to the class.

5. Electronegativity

Where do you find the most electronegative element in the periodic table?

Electronegativity is the ability of an atom in a molecule to attract the shared electrons in the chemical bond. The American chemist Linus Pauling (1901-1994) developed the most widely used scale of electronegativity values based on bond strength. The Pauling scale ranges from 0.7 to 4.0. Fluorine, the most electronegative element, is assigned a value of 4.0, and the least electronegative element, cesium, has an electronegativity value of 0.7. The electronegativity values for all the rest elements lie between these extremes.

The electronegativity of an atom is related to its ionization energy and electron affinity.

An atom with high ionization energy and high electron affinity also tends to have a high electronegativity value because of its strong attraction for electrons in a chemical bond.

Activity 2.10

Form a group and perform the following task:

Arrange each set of the given elements in order of decreasing electronegativity and explain the observed trend.

a Ba, Mg, Be, Ca

b C, Pb, Ge, Si

c Cl, F, I, Br

Present your findings to the class.

6. Metallic Character

In which region of the periodic table do you find metals and non-metals?

Metals have the tendency to lose electrons and form positive ions. As a result, metals are called electropositive elements.

In moving down a group, atomic size increases progressively, and it becomes easier for elements to lose their valence electrons and form positive ions. Therefore, metallic character increases down a group.

In the periodic table, metals and non-metals are separated by a stair step diagonal line, and elements near this border line are called metalloids. Metals are found on the left side of the line and nonmetals on its right side.

Activity 2.11

Form a group, perform the following task and present your findings to the class:

1. The members of group IVA of the periodic table are: Ge, Sn, C, Pb and Si.

Classify these elements into metals, non-metals and metalloids.

2. Explain the differences between silicon and lead in terms of:

a Atomic size

b Ionization energy

c Electron affinity

d Electronegativity

Exercise 2.4

Part I: *Choose the correct answer from the given choices*

1. Which of the following properties of the elements remain unchanged down a group?

   a Ionization energy
   b Nuclear charge
   c Electron affinity
   d Valence electrons

2. Which of the following elements has the largest atomic size?

   a Be
   b Ba
   c Ca
   d Mg

3. Which of the following elements has the lowest electronegativty?

   a F b Br c I d Cl

Part II: *Give short answers for the following*

4. What is screening effect? How does it relate to effective nuclear charge?
5. What is the relationship between first ionization energy and the metallic properties of elements?

2.3.2 Periodic Properties within a Period

Activity 2.12

Form a group and draw a rough sketch of the periodic table (*no details are required*). Based on this, perform the following activities and present your findings to the class.

1. Indicate the regions where metals, non-metals, and metalloids are located in the periodic table. (*Use different colours*).
2. Where do you find the most active metal and the most active non-metal? .

As we move from left to right across a period, the number of valence electrons of the elements increases. But the number of energy levels or main shells remains the same in a given period and electrons are filling the same energy level until stable noble gas configuration is achieved. The additional electrons from Li (2, 1) to Ne (2, 8) are added to the second shell. In fact, the period number equals the number of energy level being filled.

Table 2.7 Electron configuration and number of shells for period-2 elements.

Elements	Li	Be	B	C	N	O	F	Ne
Atomic number	3	4	5	6	7	8	9	10
Electron configuration	2, 1	2, 2	2, 3	2, 4	2, 5	2, 6	2, 7	2, 8
Number of shells (period number)	2	2	2	2	2	2	2	2

Let us now consider ***periodic properties of elements across a period.***

1. **Atomic Size:** From left to right in a given period, nuclear charge or atomic number progressively increases by one for every succeeding element. However, increasing number of valence electrons is being added to the same shell. This results in an increase in effective nuclear charge.

Activity 2.13

Form a group and perform the following task. Present your findings to the class.

The following values are given for atomic radii (in Å) of period 2 elements:

0.77, 1.34, 0.69, 0.75, 0.73, 0.82, 0.90, and 0.71.

Based on the information given;

1. fill the table given below with the appropriate values for atomic radii of the elements.
2. explain the reason for the observed trend.

Element	Li	Be	B	C	N	O	F	Ne
Atomic radii (Å)								

2. **Ionization Energy:** Two factors account for the general trend in ionization energy across a given period. These are nuclear charge and atomic size.

Activity 2.14

Form a group and perform the following task, and present your findings to the class.

The following table lists ionization energy values of period 3 elements

Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Ionization energy (kJ/mol)	496	738	578	787	1012	1000	1251	1521

1. Explain why there is a general increase in the first ionization energy across the period from Na to Ar.
2. Explain why the first ionization energy of:
   i) aluminium is lower than that of magnesium.
   ii) sulphur is lower than that of phosphorus.
   (Hint: Use sublevel configurations for the elements).

3. **Electron Affinity:** The variation in electron affinity of elements in the same period is due to changes in nuclear charge and atomic size of the elements.

   A cross a period from left to right, electron affinity increases due to an increase in effective nuclear charge. The elements show a greater attraction for an extra added electron.

4. **Electronegativity:** Across a period, a gradual change in nuclear charge and atomic size determine the trends in the electronegativity of the elements.

Activity 2.15

Form a group and perform the following task. Present your findings to the class.

The following values are given for electronegativity of period 3 elements:

2.1, 0.9, 1.5, 3.0, 1.8, 2.5 and 1.2

Based on the information given;

1. Draw a table of period 3 elements and fill with the appropriate electronegativity values corresponding to the symbols of the elements.
2. Explain the reason for the observed trend.

5. **Metallic character:** From left to right in a period, metallic character of the elements decreases. Elements on the left end of a period have a higher tendency to form positive ions. Those at the right end have a greater tendency to form negative ions. In any period, elements on the left side are metals and those on the right side are nonmetals.

   Consider the following period-3 elements and observe how the elements become more non-metallic on moving from left to right in the periodic table:

Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Character	Metals			Metalloid	Non-metals			

Metallic character decreases

Project work

Model of a periodic table

The following table shows a section of a periodic table. It is incomplete. According to the data given in table 2.8,

a Construct a model of this section of the periodic table using locally available materials to show the trends in atomic radii of the elements.

b Complete the section of the periodic table with the elements in the appropriate position.

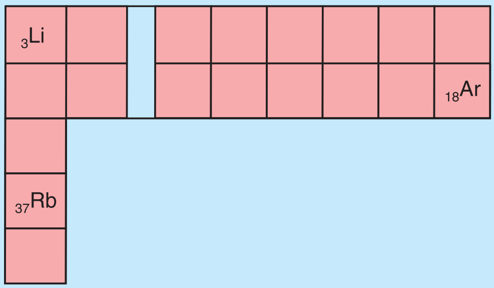


Table 2.8 Atomic number and atomic radii of some selected elements.

Element	Atomic number	Atomic radii (nm)	Element	Atomic number	Atomic radii (nm)
Li	3	0.123	Al	13	0.125
Be	4	0.089	Si	14	0.117
B	5	0.080	P	15	0.110
C	6	0.077	S	16	0.104
N	7	0.074	Cl	17	0.099
O	8	0.074	Ar	18	0.099
F	9	0.072	K	19	0.203
Ne	10	-	Rb	37	0.216
Na	11	0.157	Cs	55	0.235
Mg	12	0.136			

(*Hint: Be innovative. Use cheap local material available to you. Your teacher will help you whenever necessary.*)

Questions:

Give reasons for the increase or decrease in atomic radii that you observe in your model:

i from lithium to fluorine and sodium to chlorine

ii from lithium to cesium

Present your findings to your class teacher.

Exercise 2.5

Give appropriate answers for the following questions.

1. The element with the electron configuration of $1s^2\,2s^22p^6\,3s^23p^6$ is a:
   a metal b non-metal c metalloid d noble gas
2. Which of the main groups in the periodic table has the elements with the most negative value of electron affinity?
   a halogens b noble gases c Alkali metals d none
3. Which of the following is the correct increasing order of atomic size for the elements: Al, Ar, Na, Si?
   a Si, Na, Al, Ar c Ar, Si, Al, Na
   b Na, Al, Si, Ar d Al, Si, Ar, Na

Critical Thinking

4. Explain why the first ionization energy of nitrogen is the highest as compared to that of carbon and oxygen.
5. Compare the elements fluorine and chlorine with respect to their electronegativity and electron affinity values.

2.4 ADVANTAGES OF PERIODIC CLASSIFICATION

Competencies

By the end of this section, you will be able to:

- describe the advantages of periodic classification in studying chemistry.

Activity 2.16

Form a group and discuss the importance of the periodic table for predicting the atomic size and ionization energy of the elements. Share your idea with the class.

Why do we need the classification of elements?

The main advantages of using the periodic table are:

1. The periodic table is useful for predicting the formulas of compounds. The elements in a given group form compounds with the same atomic ratio

because of their similar electron configuration. For example, if the chemical formula of sodium oxide is Na_2O, then we can predict the formulas of the other oxides of alkali metals. These are Li_2O, K_2O, Rb_2O, and Cs_2O.

2. The periodic table is useful for predicting the physical and chemical properties of elements. For example, radium is a rare and radioactive element and therefore difficult to handle in many experiments. Since its properties can be predicted from the general trends of group IIA elements, sometimes we do not need to analyze it directly.
3. The periodic table is also useful for predicting the behaviour of many compounds. For example, oxides of the elements become more acidic across a period and more basic in character down a group. The trends in the oxides of period-3 elements vary from strongly basic oxides to amphoteric and then acidic oxides as we move across a period.

Table 2.9 Oxides of period 3 elements.

Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Basic oxide		Amphoteric oxide	Acidic oxide			

For a given element, the important information indicated below, in i, ii and iii can be read, deduced or stated from the periodic table.

(*i*) Read the
- Name
- Symbol
- Atomic number
- Atomic mass

(*ii*) Deduce the
- Number of protons and electrons
- Electron configuration (number of shells and of valence electrons)
- Character (behavior) as metal, non metal, or metalloid
- Nature (property) of the oxides to form acids, bases ...

Based on the information in i and ii, we can:

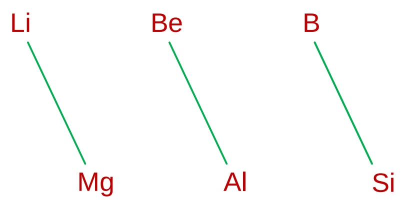
(*iii*) State the
- Period number
- Group number and
- Block type (*s*, *p*, *d* or *f*)

Diagonal Relationship

How do you account for the unexpected resemblance of the properties of the following sets of elements: Li and Mg, Be and Al, and B and Si?

In addition to the group and period relationships, the elements of s and *p* block also exhibit diagonal relationship. On moving diagonally across the periodic table, the elements show certain similarities. Though clear upon examination, these diagonal relationships are far less pronounced than the similarities within a group.

Diagonal relationship is particularly noticeable in the elements of second and third periods of the periodic table. The following illustrations show the diagonal relationship between Li and Mg, Be and Al, and B and Si.



Check List

Key terms of the unit

- *Actinide series*
- *Atomic size*
- *Block of elements*
- *Dobereiner's traids*
- *Effective nuclear charge*
- *Electron affinity*
- *Electron configuration*
- *Electronegativity*
- *Groups*
- *Inner transition elements*
- *Ionization energy*
- *Lanthanide series*
- *Law of octave*
- *Mendeleev's law*
- *Metalloids*
- *Metals*
- *Non-metals*
- *Nuclear charge*
- *Periodic law*
- *Periodic table*
- *Periods*
- *Representative elements*
- *Shielding (screening) effect*
- *Transition elements*

Unit Summary

- Early attempts of classification of the elements were made by Dobereiner and Newlands.
- Mendeleev's law states that properties of the elements are periodic functions of their atomic masses.
- Mendeleev arranged the elements based on increasing atomic mass.
- Modern periodic law states that the properties of the elements are periodic function of their atomic numbers.
- Periods are horizontal rows, and groups are vertical columns of the elements in the periodic table.
- Elements in the same group show similar chemical properties.
- Elements are classified as representative, transition, and rare-earth elements. This classification is based on the type of sub-level (*s*, *p*, *d*, or *f*) being filled.
- Ionization energy is the energy required to remove the outermost shell electron from an isolated gaseous atom.
- Electron affinity of an element is the energy released when an electron is added to an isolated gaseous atom to form a gaseous ion.
- Metallic character is the tendency to lose electrons and form positively charged ions.
- Electronegativity of an element is its ability to attract electrons.
- Trends in atomic size determine the trends in ionization energy, electron affinity, electronegativity, and metallic character of the elements in the periodic table.
- Atomic size itself is determined by the number of energy levels, nuclear charge, and effective nuclear charge.
- Periodic properties of elements, such as ionization energy, electron affinity, electronegativity, etc. show regular variation within a group or period.

REVIEW EXERCISE ON UNIT 2

Part I: Identify whether each of the following statements is true or false. Give your reasons when you consider a statement to be false.

1. Metallic properties of the elements increase from left to right within a period.
2. Elements in a group have consecutive atomic numbers.
3. All elements with high ionization energy also have high electron affinity.

4. All the elements that belong to s and p-blocks are metals.
5. The modern periodic law was proposed by Mosley.
6. As the atomic number of elements increases in the periodic table, their atomic radius also increases.
7. Transition metals are found in four periods. Each corresponds to the filling of valence electrons in the 3*d*, 4*d*, 5*d*, and 6*d* orbitals.

Part II: Multiple Choices Type Questions

8. The element with atomic number 35 belongs to:
   a s-block c d-block
   b p-block d f-block
9. An element with the valence electron configuration $3s^2 3p^2$ belongs to group:
   a IIIA c VIA
   b IVA d VIIA
10. An element with the valence electron configuration $3s^2 3p^6$ belongs to period:
    a 2 c 8
    b 6 d 3
11. An element with the valence electron configuration $2s^2 2p^4$ belongs to:
    a alkali metals c halogens
    b oxygen group d noble gases
12. The property of the element with atomic number 18 resembles that of the element with atomic number:
    a 8 b 36 c 19 d 40
13. An element with high ionization energy and high electron affinity also tends to have a high value of:
    a electronegativity c atomic size
    b electropositivity d metallic character

Part III: To which group, period, and sublevel block do the following elements belong?

	Group	Period	Block
14. Calcium	______	______	______
15. Sulphur	______	______	______
16. Iron	______	______	______
17. Zinc	______	______	______
18. Xenon	______	______	______

Part IV: Give short answers

19. Explain, with examples, the meaning of diagonal relationship.
20. From the given elements : Na, P, Ca, and Br; which one has the:
    a highest first ionization energy?
    b smallest atomic size?
    c most metallic character?
21. On the basis of electron configuration, explain why nitrogen is less electronegative than oxygen?
22. On the basis of electron configuration, explain why atomic size of sodium is larger than that of chlorine?