

UNIT 1 THE PERIODIC TABLE

Structure

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1.1 INTRODUCTION

Scientists from the very beginning have attempted to systematise the knowledge they gain through their observations and experiments. Development of the periodic law and the periodic table of the elements is one such attempt. This has brought order in the study of the vast chemistry of more than a hundred elements known now. Therefore, it is quite natural to begin your study of inorganic chemistry with the study of the periodic table in this unit.

By the mid-nineteenth century, more than 60 elements were known and many more were being discovered. The rate of discovery of the new elements was so fast that the chemists started wondering "where would it all lead to"? "Has nature provided a limit to the number of elements and if so, how would one know about it"? During this period, it was also realised that certain groups of elements exhibited similar physical and chemical properties. Was it a mere coincidence or did a relationship exist among the properties of the elements? Attempts to reply such probing questions ultimately resulted in the development of the periodic law.

Objectives

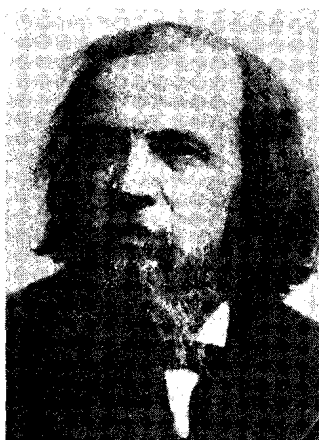
After studying this unit, you should be able to:

- explain the origin of various forms of the periodic table,
- state and explain the periodic law,
- define periodicity,
- explain the cause of periodicity in properties,
- classify the elements into four categories according to their electronic configuration, and
- write the IUPAC symbols of the elements of atomic number greater than 100.

1.2 EARLIER ATTEMPTS AT CLASSIFICATION

One of the earliest attempts to classify elements was to divide them into metals and non-metals. As you know metallic elements have a certain lustrous appearance, are malleable (can be beaten into thin sheets) and ductile (can be drawn into wires), conduct heat and electricity, and form basic oxides. Non-metallic elements, on the other hand, do not have any one characteristic appearance, but they are brittle, do not generally conduct heat and electricity, and form acidic oxides. As more elements were discovered and knowledge of physical and chemical properties was refined, it became clear that within these two divisions of elements, there existed families of elements whose properties varied systematically from each other. Furthermore, certain elements, the metalloids, possessed properties intermediate between the two divisions. Therefore, attempts were made to search for other classifications.

In 1829, **J. W. Dobereiner** observed that there existed certain groups of three elements which he called **triads**. He also observed that elements in a triad not only had similar properties but also the atomic weight of the middle element was approximately an average of the atomic weights of the other two elements of the triad. A few examples cited by him were: Li, Na, K; Ca, Sr, Ba; S, Se, Te and Cl, Br, I. Although, Dobereiner's relationship seemed to work only for a few elements, he was the first to point out a systematic relationship among the elements.



Dmitri Ivanovich Mendeleev, 1834-1907. A Russian Chemist and a Professor of Chemistry at the Technological Institute of St. Petersburg (now Leningrad). He developed the periodic table of the elements and wrote an outstanding Chemistry textbook. In 1906, he failed by one vote to receive the Nobel Prize in Chemistry.

The next reported attempt to classify elements was made by a French geologist, **A. de Chancourtois** in 1862. He arranged the then known elements in order of increasing atomic weight on a line which spiralled around a cylinder from bottom to top. He divided the circumference of the cylinder into 16 sub-divisions and pointed out that the elements with similar properties ended up above each other on adjacent turns of the spiral. He called his representation as a '**Telluric Screw**'. This also did not attract much attention.

In 1864, the English chemist **John Newlands** reported his "**Law of Octaves**". He suggested that if the elements be arranged in the order of increasing atomic weight, every eighth element had properties similar to the first element just like every eighth note resembles the first in octaves of music. For example, he arranged the elements in the following manner:

Element	Li	Be	B	C	N	O	F
At. Wt.	7	9	11	12	14	16	19
Element	Na	Mg	Al	Si	P	S	Cl
At. Wt.	23	24	27	29	31	32	35.5
Element	K	Ca	Ti	Cr			
At. Wt.	39	40	48	52			

Thus we see, K resembles Na and Li; Ca resembles Mg and Be; Al resembles B; Si resembles C, etc. Newlands' law of octaves was rejected due to two reasons. Firstly, it did not hold good for elements heavier than Ca. Secondly, he believed that there existed some mystical connection between music and chemistry.

The periodic law, as we know it today owes its development to two chemists, **Lothar Meyer** and **Dmitri Mendeleev**, each one working independently propounded it simultaneously in 1869. Lothar Meyer reported that when physical properties like atomic volume, boiling point, etc., were plotted against atomic weight, a periodically repeating curve was obtained in each case. Fig. 1.1 is a graph showing the variation in atomic volume with the atomic number (similar curve was obtained by Lothar Meyer by plotting atomic volume versus atomic weight). The atomic volume behaviour is periodic. It goes through cycles, dropping from a sharp maximum to a minimum and then sharply rising again. Each of the cycles is called a **period**. The location of elements on the peaks or in the troughs has an important correlation with their chemical reactivity. The elements at the peaks (e.g., alkali metals) are the most reactive, those in the troughs (e.g., noble metals) are characteristically less reactive.

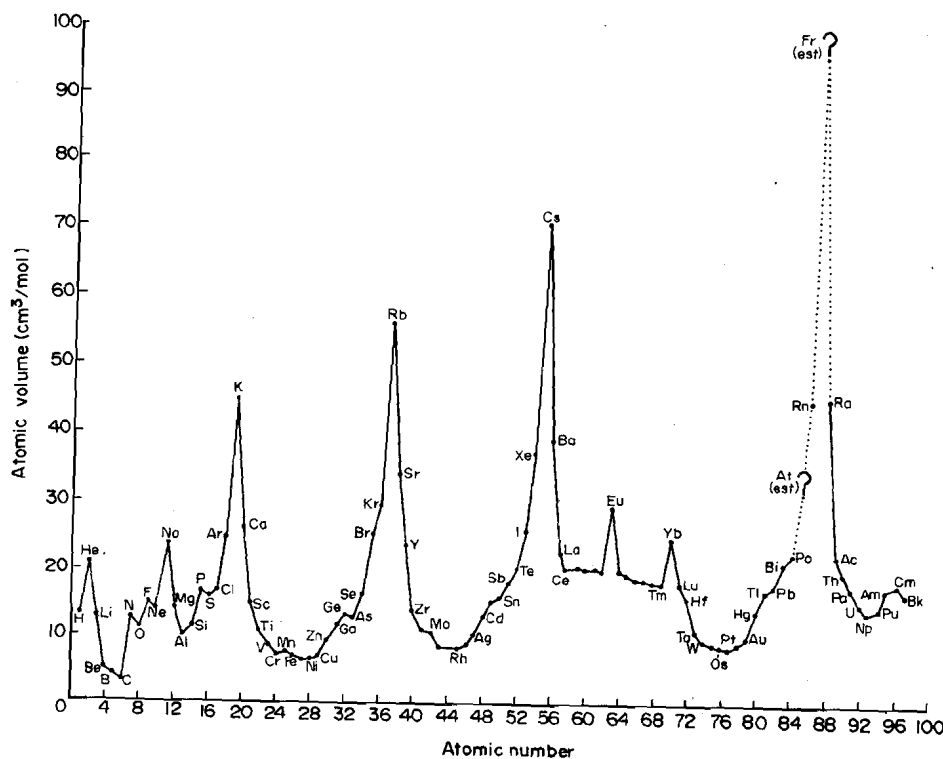


Fig. 1.1: Periodic dependence of atomic volume on atomic number.

Mendeleev's Periodic Law

Mendeleev, on the other hand, used chemical properties like valence and formulae of hydrides, chlorides and oxides of the elements to illustrate his periodic law. According to Mendeleev's periodic law, "if the elements are arranged sequentially in the order of increasing atomic weight, a periodic repetition, i.e., periodicity in properties is observed." Mendeleev arranged elements in horizontal rows and vertical columns in order of increasing atomic weight so that the elements having similar properties were kept in the same vertical column. Fig. 1.2 shows the periodic table devised by Mendeleev in 1871.

Series	Group I — R ₂ O	Group II — RO	Group III — R ₂ O ₃	Group IV RH ₄ RO ₂	Group V RH ₃ R ₂ O ₅	Group VI RH ₂ RO ₃	Group VII RH R ₂ O ₇	Group VIII — RO ₄
1	H=1							
2	Li=7	Be=9.4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27.3	Si=28	P=31	S=32	Cl=35.5	
4	K=39	Ca=40	=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59 Ni=59, Cu=63
5	(Cu=63)	Zn=65	=68	=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	=100	Ru=104, Rh=104 Pd=106, Ag=108
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	I=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	—
9	—	—	—	—	—	—	—	—
10	—	—	?Er=178	?La=180	Ta=182	W=184		Os=195, Ir=195 Pt=198, Au=199
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208			
12	—	—		Th=231	—	U=240	—	—

Fig. 1.2: Mendeleev's periodic table of 1871; against each element is the 1871 value of atomic weight

Although Newlands and Lothar Meyer also contributed in developing the periodic law, the main credit goes to Mendeleev because of the following reasons:

- He included, along with his table, a detailed analysis of the properties of all the known elements and correlated a broad range of physical and chemical properties with atomic weights.
- He kept his primary goal of arranging similar elements in the same group, quite clearly. Therefore, he was bold enough in reversing the order of certain elements. For example, iodine with lower atomic weight than that of tellurium (Group VI) was placed in Group VII along with fluorine, chlorine and bromine because of similarities in properties.
- He also corrected the atomic weights of certain elements to include them in proper groups. For example, he corrected the atomic weights of beryllium (from 13.5 to 9), and indium (from 76 to 114) without doing any actual measurements; his conjecture was proved correct as Be and In with equivalent weights of 4.5 and 38, respectively are actually bivalent and trivalent.
- Keeping to his primary goal of arranging similar elements in the same vertical column (group), he realised that some of the elements were still undiscovered and therefore, left their places vacant in the table and predicted their properties. He predicted the existence in nature of over ten new elements and predicted properties of three of them, e.g., eka-boron (scandium), eka-aluminium (gallium) and eka-silicon (germanium) from the properties of known elements surrounding them. When these elements were discovered, Mendeleev's predictions proved to be amazingly accurate. This you can see by comparing the predicted and observed properties of eka-aluminium (gallium) and eka-silicon (germanium) given in Table 1.1.

It is very interesting to know that out of more than ten elements predicted by Mendeleev, eka-aluminium was discovered first of all and named gallium by **Lecoq de Boisbaudran** in 1875. He determined the density and melting point of gallium and also studied its compounds. He first reported the density of gallium as $4.7 \times 10^3 \text{ kg m}^{-3}$. But when he received a letter from Mendeleev saying that everything was quite correct in his description of gallium except for the density which should be $6.0 \times 10^3 \text{ kg m}^{-3}$, Boisbaudran was highly

Table 1.1 Comparison of predicted and observed properties of eka-aluminium (gallium) and eka-silicon (germanium)

Property	Predicted by Mendeleev for eka-aluminium	Observed for gallium	Predicted by Mendeleev for eka-silicon	Observed for germanium
At. weight	68	69.72	72	72.59
Density (kg m^{-3})	6.0×10^3	5.9×10^3	5.5×10^3	5.3×10^3
Melting point (K)	low	302.8	High	1220 K
Reaction with acids and alkalies	slow	slow	slow	Reacts with conc. acids and alkalies
Formula of oxide	E_2O_3	Ga_2O_3	—	—
Density of oxide (kg m^{-3})	5.5×10^3	5.88×10^3	—	—
Formula of chloride	ECl_3	GaCl_3	EsCl_4	GeCl_4
Boiling point of chloride (K)	Volatile	474	373	357

surprised. He thought could it be possible that someone else also had discovered gallium before him? He determined the density of gallium again after purifying the metal more carefully. You will be equally surprised to know that Mendeleev was right; the density of gallium is $5.9 \times 10^3 \text{ kg m}^{-3}$.

Another four years later in 1879, the Swedish chemist **Lars Fredrick Nilson** discovered a new element which he named scandium (for Scandinavia). When its properties were investigated, it became absolutely obvious that it was nothing but eka-boron predicted by Mendeleev. In 1886, the German chemist **Winkler** discovered eka-silicon which he renamed as germanium (for Germany). Winkler found that germanium completely corresponds to Mendeleev's eka-silicon. You can yourself see by comparing the properties of eka-silicon and germanium given in Table 1.1.

In this way, the discovery of these three elements provided a striking proof for the validity of the periodic law. The development of periodic law is an excellent example where careful observation, critical analysis of available data without any preconceived notions and scientific foresight led to the discovery of a fundamental law of nature. Thus, when Mendeleev arranged elements in order of increasing atomic weight and critically analysed the properties of then known elements, he discovered that the properties of any element are an average of the properties of its neighbours in the periodic table. On this basis he predicted the properties of undiscovered elements representing the gaps in his table.

SAQ 1

a) What property did Mendeleev use to classify the elements in his periodic table?

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b) Enumerate four defects in the Mendeleev's periodic table.

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c) For appreciating the excellent achievement of Mendeleev and for assessing your own understanding of the Periodic Law, imagine for a while that one of the elements, say calcium, is still undiscovered. Predict its properties from the properties of the known surrounding elements and then, compare your predictions with those already known.

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1.3 MODERN PERIODIC LAW

You must have studied in previous classes that there were some anomalies in Mendeleev's original periodic table. Between 1869 - 1907 he, therefore, tried to improve it further. However, the most significant improvement of his periodic table came through the discovery of the concept of atomic number in 1913 by **Henry Moseley**, who suggested that the atomic number of an element is a more fundamental property than its atomic weight. Mendeleev's Periodic Law was, therefore, accordingly modified. This is now known as **the Modern Periodic Law** and can be stated as "**the properties of elements are periodic functions of their atomic numbers**". Mendeleev could not arrive at this conclusion because atomic numbers were not known at that time. However, the table given by him (Fig. 1.2) broadly resembles the modern table (Fig. 1.3) because the atomic weight and atomic number vary almost in a parallel manner.



Henry G.J. Moseley, 1887-1915, A brilliant English physicist discovered the concept of atomic number in 1913. He had to go to battlefield during World War I and got killed in 1915.

In the modern plan of classification (Fig. 1.3), elements are arranged in the table in order of their increasing atomic numbers, and this removes most of the anomalies of Mendeleev's

Per-iods	A I B						A VII B		A VIII B		
	1	1 H 1.008						1 H 1.008	2 He 4.003		
2	3 Li 6.941	4 Be 9.012	5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18			
3	11 Na 22.99	12 Mg 24.31	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95			
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.80	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.71	
	29 Cu 63.54	30 Zn 65.37	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.91	36 Kr 83.80			
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.4	
	47 Ag 107.87	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.30			
6	55 Cs 132.91	56 Ba 137.54	57 La* 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	
	79 Au 197.97	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.98	84 Po 210	85 At 210	86 Rn 222			
7	87 Fr 223	88 Ra 226.03	89 Ac** 227.03	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une		

* LANTHANIDES

58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 146.92	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
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** ACTINIDES

90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu 239.05	95 Am 241.06	96 Cm 247.07	97 Bk 249.08	98 Cf 251.08	99 Es 254.09	100 Fm 257.10	101 Md 258.10	102 No 255.0	103 Lr 257.0
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Fig. 1.3: Modern periodic table in the form devised by Mendeleev

original periodic table. The positions of K and Ar, Co and Ni, Te and I do not remain anomalous any longer. There was no place for isotopes of an element in Mendeleev's original periodic table. As isotopes of an element have the same atomic number, they can all be placed at one and the same place in the periodic table. You also know the atomic number cannot be fractional and it increases by one integer from one element to the next; it has thus placed a limit on the number of elements. Thus today, 109 elements (from 1 to 109) have been discovered and any more elements that may be discovered in future will be beyond 109.

Long Form of the Periodic Table

You have now learnt that in the modern form of Mendeleev's periodic table, elements are arranged in seven horizontal rows and eight vertical columns. Normal and transition elements belonging to A and B subgroups of a group were placed in one and the same column of the table. In the long form of the periodic table (Fig. 1.4), elements are arranged in eighteen vertical columns by keeping the elements belonging to A and B subgroups in separate columns. You have also studied that the Group VIII B of Mendeleev's periodic table contains three triads — Fe, Co, Ni (4th period); Ru, Rh, Pd (5th period) and Os, Ir, Pt (6th period). But in the long form of the table, each element of the triad is kept in a separate column. So, the group VIII B occupies three columns of the table. Thus, you can see that the long form of periodic table is an extension of the modern periodic table. It is also related to the electronic configuration of elements and displays the order in which the electrons are filled in *s*, *p*, *d* and *f* orbitals about which you will study in Section 1.5.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
IA	IIA	IIIB	IVB	VB	VIB	VII B		VIII B		IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
←-sBlock->												←-p-block elements->					
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31	←-d-block elements->										13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.90	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.71	29 Cu 63.54	30 Zn 65.37	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.91	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.6	47 Ag 107.87	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.30
55 Cs 132.91	56 Ba 137.34	57 La* 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 197.97	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.98	84 Po 210	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226.03	89 Ac** 227.03	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une									
←-f-block elements->																	
* Lanthanides		58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 146.92	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.93	71 Lu 174.97		
** Actinides		90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu 239.05	95 Am 241.06	96 Cm 247.07	97 Bk 249.08	98 Cf 251.08	99 Es 254.09	100 Fm 257.10	101 Md 258.10	102 No 255	103 Lr 257		

Fig. 1.4 : Long form of the periodic table

Originally Mendeleev gave A and B designations to the groups containing normal and transition elements, respectively. However, in his periodic table, this division into A and B groups is often done arbitrarily. In different books, for the elements of III to VIII groups, this designation of A and B groups is often reversed. To avoid this controversy,

International Union of Pure and Applied Chemistry (IUPAC) has adopted Arabic Numerals 1, 2, 3, ..., 18 as the newest group designations in the long form of periodic table. In this system, therefore, the alkali and alkaline earth metals constitute Group 1 and 2; transition elements of Sc to Zn families become Groups 3, 4, 5, ..., 12 and finally the *p*-block elements become Groups 13, 14, ..., 18 of the table. In this unit and throughout the rest of this course we will use this IUPAC representation for the elements of different groups in the periodic table.

SAQ 2

- a) Which group of elements appears in the modern periodic table, but did not appear in Mendeleev's original table? Why?

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- b) What is the relationship between the atomic number and the periodic classification of elements?

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1.4 NOMENCLATURE OF ELEMENTS HAVING $Z > 100$

It has been a historical practice to allow the discoverer of an element to assign the element's name. In recent times, this had led to some controversy, because elements with very high atomic numbers are so unstable that only minute quantities of them, sometimes only one or two atoms, are prepared before scientists claim credit for their discovery. This has led to questions of the reliability of the data and whether the said new element has in fact been discovered. For example, both American and Soviet scientists claimed credit for discovering element 104. The Americans named it rutherfordium and the Soviets named it kurchatovium. To avoid this problem, the IUPAC has made an official recommendation that until a new element's discovery has been proved, a systematic nomenclature be applied according to the following IUPAC nomenclature rules:

- The name be derived directly from the atomic number of element using the following numerical roots:

0	1	2	3	4	5	6	7	8	9
nil	un	bi	tri	quad	pent	hex	sept	oct	enn

- The roots be put together in the order of the digits which make up the atomic number and be terminated by 'ium', an ending occurring in the names of all the metallic elements, as these are. The final 'n' of enn be dropped when it occurs before 'nil' and the final 'i' of 'bi' and 'tri' be dropped when it occurs before 'ium'.
- The symbol of the element be composed of the initial letters of the numerical roots which make up the name.

In Table 1.2 systematic names and symbols of elements having $Z=101$ to 106 derived by IUPAC norms are listed.

Table 1.2: Systematic names for elements having $Z > 100$.

Atomic Number	Systematic Name	Symbol	Trivial Name
101	unnilunium	Unu	Mendelevium
102	unnilbium	Unb	Nobelium
103	unniltrium	Unt	Lawrencium
104	unnilquadium	Unq	—
105	unnilpentium	Unp	—
106	unnihexium	Unh	—

SAQ 3

Write the IUPAC names and symbols for elements having $Z = 107$ to 114.

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1.5 ELECTRONIC CONFIGURATION OF ATOMS

In the preceding section you have studied the arrangement of elements in the long form of the periodic table. In this section you will study the electronic configurations of atoms to understand the cause of periodicity in the properties of the elements. The electronic configurations of isolated atoms of elements are verified experimentally by a detailed analysis of atomic spectra which are too complex for discussion here. However, the electronic configurations of the atoms can be predicted with the help of *aufbau* or the building up process.

In the *aufbau* process, it is assumed that there exists a set of empty hydrogen like orbitals around the nucleus of an atom. The electronic configuration of the atom in the ground state is then derived by adding electrons one at a time to the orbitals of lowest energy in the sequence shown by arrows in Fig. 1.5 given below.

Order of orbitals in the above filling diagram you can remember in the form of $n+l$ rule. According to this rule, in the building up of electronic configuration of the elements, the subshell with the lowest value of $n+l$ fills first; when two subshells have the same value of $n+l$, the orbital with lower value of n fills first. This useful rule reminds us that the energy of subshells of multi-electron atoms, depends upon the value of both the quantum numbers n and l , but mainly on the value of n . For example, which fills first, the $5s$ or the $4p$ subshell? For the $5s$ subshell the value of $n+l = 5 + 0 = 5$; for $4p$ subshell also, the value of $n+l = 4 + 1 = 5$, but the $4p$ subshell has the lower value of the principal quantum number n and therefore, it fills first.

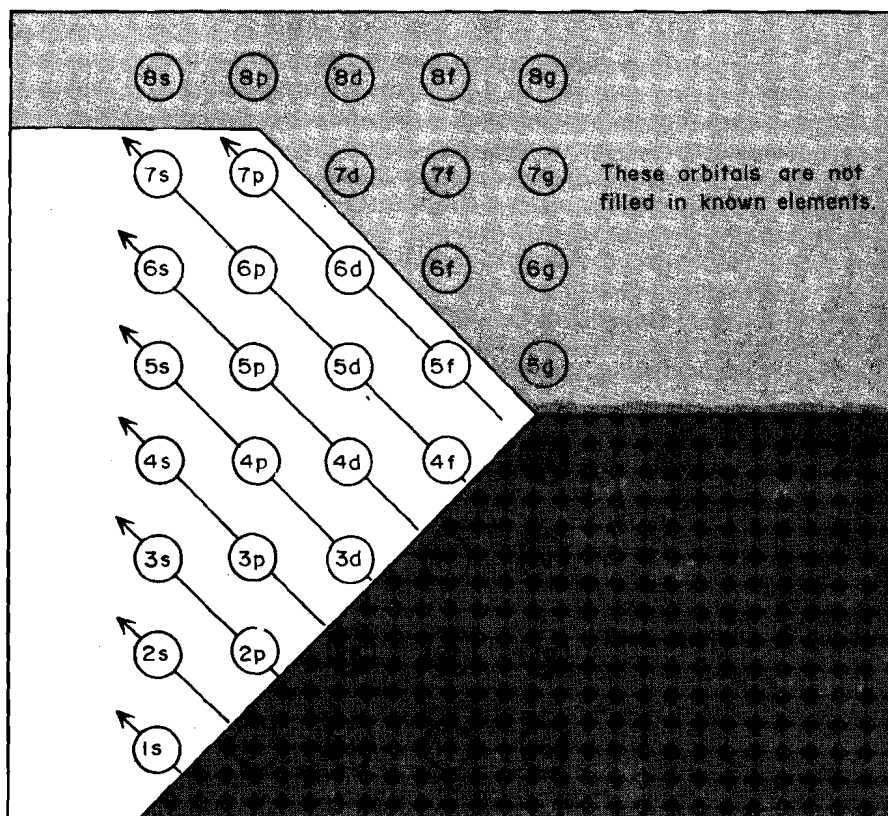


Fig. 1.5: Order of filling of atomic orbitals in polyelectronic atoms

Filling of electrons in orbitals is governed by **Pauli's Exclusion Principle** and **Hund's Rule**. According to the **Exclusion Principle** no two electrons in the same atom can have the same values of all the four quantum numbers. Even if two electrons have the same values of n , l and m_l , they will differ in their m_s values. In other words, an orbital can have at the most two electrons of opposite spin. Since there is only one s -orbital for any given value of n , it can contain only two electrons. However, the three p orbitals, for any given value of n , can contain six electrons, the five d orbitals can hold a total of ten electrons and the seven f orbitals can have fourteen electrons. Permitted combinations of all the four quantum numbers for the electrons in different orbitals are given in Table 1.3.

Table 1.3: Permitted combinations of quantum numbers for s , p , d and f orbitals.

n	l	m_l	m_s	Common name	Number of electrons
1	0	0	$\pm 1/2$	1s	2
2	0	0	$\pm 1/2$	2s	2
2	1	-1	$\pm 1/2$	2p	6
		0	$\pm 1/2$		
		+1	$\pm 1/2$		
3	0	0	$\pm 1/2$	3s	2
3	1	-1	$\pm 1/2$	3p	6
		0	$\pm 1/2$		
		+1	$\pm 1/2$		
3	2	-2	$\pm 1/2$	3d	10
		-1	$\pm 1/2$		
		0	$\pm 1/2$		
		+1	$\pm 1/2$		
		+2	$\pm 1/2$		
4	0	0	$\pm 1/2$	4s	2
4	1	-1	$\pm 1/2$	4p	6
		0	$\pm 1/2$		
		+1	$\pm 1/2$		
4	2	-2	$\pm 1/2$	4d	10
		-1	$\pm 1/2$		
		0	$\pm 1/2$		
		+1	$\pm 1/2$		
		+2	$\pm 1/2$		
4	3	-3	$\pm 1/2$	4f	14
		-2	$\pm 1/2$		
		-1	$\pm 1/2$		
		0	$\pm 1/2$		
		+1	$\pm 1/2$		
		+2	$\pm 1/2$		
		+3	$\pm 1/2$		
5	0		etc.		

Hund's rule of maximum multiplicity states that as far as possible in a given atom in the ground state, electrons in the same subshell will occupy different orbitals and will have parallel spins. That is when electrons are added to orbitals of the same energy such as three p orbitals or five d orbitals, one electron will enter each of the available orbitals before a second electron enters any one orbital. This behaviour, can be explained in terms of electron-electron repulsion. Two electrons in separate orbitals feel less repulsion than two electrons paired in the same orbital. For example, carbon in the ground state, has the configuration $1s^2 2s^2 2p_x^1 2p_y^1$ rather than $1s^2 2s^2 2p_z^2$.

So far you have studied the rules governing the filling of electrons in the orbitals of atoms. We shall now consider the electronic configurations of all the elements in the periodic table. These are given in Table 1.4.

Period 1: This is the smallest of all the periods of the table. Hydrogen ($Z=1$) and helium ($Z=2$) are the two elements belonging to this period. The electronic configurations of hydrogen and helium are $1s^1$ and $1s^2$, respectively. Thus, the $1s$ orbital which is the only orbital corresponding to $n=1$ is completely filled. The $1s^2$ configuration of helium is also represented by [He].

Period 2: This period contains elements from lithium ($Z=3$) to neon ($Z=10$). In lithium and beryllium, the filling of $2s$ orbital takes place. Then in the next six elements from boron to neon, the $2p$ orbitals are filled. Neon thus has the electronic configuration [He] $2s^2 2p^6$ which is also represented by [Ne]. At this stage the shell having $n=2$ is complete.

Table 1.4: Ground-state electronic configuration of gaseous atoms

Z	Symbol	Configuration as [Core] plus "Outermost" Orbitals
1	H	1s ¹
2	He	1s ² , or [He]
3	Li	[He]2s ¹
4	Be	[He]2s ²
5	B	[He]2s ² 2p ¹
6	C	[He]2s ² 2p ²
7	N	[He]2s ² 2p ³
8	O	[He]2s ² 2p ⁴
9	F	[He]2s ² 2p ⁵
10	Ne	[He]2s ² 2p ⁶ , or [Ne]
11	Na	[Ne]3s ¹
12	Mg	[Ne]3s ²
13	Al	[Ne]3s ² 3p ¹
14	Si	[Ne]3s ² 3p ²
15	P	[Ne]3s ² 3p ³
16	S	[Ne]3s ² 3p ⁴
17	Cl	[Ne]3s ² 3p ⁵
18	Ar	[Ne]3s ² 3p ⁶ , or [Ar]
19	K	[Ar]4s ¹
20	Ca	[Ar]4s ²
21	Sc	[Ar]3d ¹ 4s ²
22	Ti	[Ar]3d ² 4s ²
23	V	[Ar]3d ³ 4s ²
24	Cr	[Ar]3d ⁵ 4s ¹
25	Mn	[Ar]3d ⁵ 4s ²
26	Fe	[Ar]3d ⁶ 4s ²
27	Co	[Ar]3d ⁷ 4s ²
28	Ni	[Ar]3d ⁸ 4s ²
29	Cu	[Ar]3d ¹⁰ 4s ¹
30	Zn	[Ar]3d ¹⁰ 4s ²
31	Ga	[Ar]3d ¹⁰ 4s ² 4p ¹
32	Ge	[Ar]3d ¹⁰ 4s ² 4p ²
33	As	[Ar]3d ¹⁰ 4s ² 4p ³
34	Se	[Ar]3d ¹⁰ 4s ² 4p ⁴
35	Br	[Ar]3d ¹⁰ 4s ² 4p ⁵
36	Kr	[Ar]3d ¹⁰ 4s ² 4p ⁶ , or [Kr]
37	Rb	[Kr]5s ¹
38	Sr	[Kr]5s ²
39	Y	[Kr]4d ¹ 5s ²
40	Zr	[Kr]4d ² 5s ²
41	Nb	[Kr]4d ⁴ 5s ¹
42	Mo	[Kr]4d ⁵ 5s ¹
43	Tc	[Kr]4d ⁵ 5s ²
44	Ru	[Kr]4d ⁷ 5s ¹
45	Rh	[Kr]4d ⁸ 5s ¹
46	Pd	[Kr]4d ¹⁰
47	Ag	[Kr]4d ¹⁰ 5s ¹
48	Cd	[Kr]4d ¹⁰ 5s ²
49	In	[Kr]4d ¹⁰ 5s ² 5p ¹
50	Sn	[Kr]4d ¹⁰ 5s ² 5p ²
51	Sb	[Kr]4d ¹⁰ 5s ² 5p ³
52	Te	[Kr]4d ¹⁰ 5s ² 5p ⁴
53	I	[Kr]4d ¹⁰ 5s ² 5p ⁵
54	Xe	[Kr]4d ¹⁰ 5s ² 5p ⁶ , or [Xe]
55	Cs	[Xe]6s ¹
56	Ba	[Xe]6s ²
57	La	[Xe]5d ¹ 6s ²
58	Ce	[Xe]4f ¹ 5d ¹ 6s ²
59	Pr	[Xe]4f ³ 6s ²
60	Nd	[Xe]4f ⁴ 6s ²
61	Pm	[Xe]4f ⁵ 6s ²
62	Sm	[Xe]4f ⁶ 6s ²
63	Eu	[Xe]4f ⁷ 6s ²
64	Gd	[Xe]4f ⁷ 5d ¹ 6s ²

Z	Symbol	Configuration as [Core] plus "Outermost" Orbitals
65	Tb	[Xe]4f ⁹ 6s ²
66	Dy	[Xe]4f ¹⁰ 6s ²
67	Ho	[Xe]4f ¹¹ 6s ²
68	Er	[Xe]4f ¹² 6s ²
69	Tm	[Xe]4f ¹³ 6s ²
70	Yb	[Xe]4f ¹⁴ 6s ²
71	Lu	[Xe]4f ¹⁴ 5d ¹ 6s ²
72	Hf	[Xe]4f ¹⁴ 5d ² 6s ²
73	Ta	[Xe]4f ¹⁴ 5d ³ 6s ²
74	W	[Xe]4f ¹⁴ 5d ⁴ 6s ²
75	Re	[Xe]4f ¹⁴ 5d ⁵ 6s ²
76	Os	[Xe]4f ¹⁴ 5d ⁶ 6s ²
77	Ir	[Xe]4f ¹⁴ 5d ⁷ 6s ²
78	Pt	[Xe]4f ¹⁴ 5d ⁹ 6s ¹
79	Au	[Xe]4f ¹⁴ 5d ¹⁰ 6s ¹
80	Hg	[Xe]4f ¹⁴ 5d ¹⁰ 6s ²
81	Tl	[Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ¹
82	Pb	[Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ²
83	Bi	[Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ³
84	Po	[Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ⁴
85	At	[Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ⁵
86	Rn	[Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ⁶ , or [Rn]
87	Fr	[Rn]7s ¹
88	Ra	[Rn]7s ²
89	Ac	[Rn]6d ¹ 7s ²
90	Th	[Rn]6d ² 7s ²
91	Pa	[Rn]5f ² 6d ¹ 7s ²
92	U	[Rn]5f ³ 6d ¹ 7s ²
93	Np	[Rn]5f ⁴ 6d ¹ 7s ²
94	Pu	[Rn]5f ⁶ 7s ²
95	Am	[Rn]5f ⁷ 7s ²
96	Cm	[Rn]5f ⁷ 6d ¹ 7s ²
97	Bk	[Rn]5f ⁹ 7s ²
98	Cf	[Rn]5f ¹⁰ 7s ²
99	Es	[Rn]5f ¹¹ 7s ²
100	Fm	[Rn]5f ¹² 7s ²
101	Md	[Rn]5f ¹³ 7s ²
102	No	[Rn]5f ¹⁴ 7s ²
103	Lr	[Rn]5f ¹⁴ 6d ¹ 7s ²
104	Unq	[Rn]5f ¹⁴ 6d ² 7s ²
105	Unp	[Rn]5f ¹⁴ 6d ³ 7s ²
106	Unh	[Rn]5f ¹⁴ 6d ⁵ 7s ¹
107	Uns	[Rn]5f ¹⁴ 6d ⁵ 7s ²
108	Uno	[Rn]5f ¹⁴ 6d ⁶ 7s ²
109	Une	[Rn]5f ¹⁴ 6d ⁷ 7s ²

Period 3: Similar to period 2, this period also consists of 8 elements from sodium ($Z=11$) to argon ($Z=18$). In these elements, $3s$ and $3p$ orbitals are successively filled in the sequence analogous to that in period 2. Thus, argon has the electronic configuration $[\text{Ne}]2s^22p^6$ represented as $[\text{Ar}]$. Although, the third principal shell ($n=3$) can accommodate 10 more electrons in $3d$ orbitals but filling of $4s$ orbital takes place first because of its lower energy.

Period 4: This period contains 18 elements from potassium ($Z=19$) to krypton ($Z=36$). In K and Ca, the first two elements of this period, the successive electrons go into the $4s$ orbital giving them the configuration $[\text{Ar}]4s^1$ and $[\text{Ar}]4s^2$, respectively. Then in the following 10 elements (Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu and Zn), filling of hitherto unoccupied $3d$ orbitals takes place. Thus, the electronic configuration of zinc becomes $[\text{Ar}]3d^{10}4s^2$. Occasionally an electron from $4s$ orbital is shifted out of turn to the $3d$ orbital due to the higher stability of half filled and completely filled orbitals. For example, Cr ($Z=24$) and Cu ($Z=29$) have the configurations $[\text{Ar}]3d^54s^1$, and $[\text{Ar}]3d^{10}4s^1$ instead of expected $[\text{Ar}]3d^44s^2$ and $[\text{Ar}]3d^94s^2$, respectively. After the $3d$ level is filled, in the next six elements of this period, i.e., Ga, Ge, As, Se, Br and Kr, the $4p$ orbitals are gradually filled and Kr has the electronic configuration $[\text{Ar}]3d^{10}4s^24p^6$ represented as $[\text{Kr}]$.

Period 5: The next 18 elements from rubidium ($Z=37$) to xenon ($Z=54$) belong to this period. In the building up of the atoms of these elements $5s$, $4d$ and $5p$ orbitals are

successively filled just as the 4s, 3d and 4p orbitals are filled in the elements of period 4. In Rb (Z=37) and Sr (Z=38), the 5s orbital is filled. After that, in elements from Y (Z=39) to Cd (Z=48) filling of 4d orbitals takes place. You can see from the Table 1.4 that once again there are minor irregularities in the distribution of electrons between 4d and 5s orbitals. For example, Mo (Z=42) and Ag (Z=47) have respectively $[\text{Kr}]4d^55s^1$ and $[\text{Kr}]4d^{10}5s^1$ configurations similar to those of Cr and Cu, respectively. Anomalous electronic configurations of Nb, Ru, Rh and Pd cannot be explained in simple terms. You have to, therefore, remember them as exceptions. Now in the next six elements, i.e., In, Sn, Sb, Te, I and Xe filling of 5p orbitals takes place and thus Xe (Z=54) attains $[\text{Kr}]4d^{10}5s^25p^6$ configuration.

Period 6: This period contains 32 elements from caesium (Z = 55) to radon (Z=86) in which the 6s, 4f, 5d and 6p orbitals are filled. The first two elements of this period have configurations analogous to those of the corresponding members of the lower periods. Thus caesium (Z=55) and barium (Z=56) have $[\text{Xe}]6s^1$ and $[\text{Xe}]6s^2$ configurations, respectively. According to the *aufbau* principle, in the next element La (Z=57), the additional electron should enter 4f orbital. Instead, it goes to the 5d orbital and La has the configuration $[\text{Xe}]5d^16s^2$. But why? The extra electron in the building up of La atom goes to 5d orbital instead of 4f orbital because in La atom, the 5d and 4f orbitals have almost the same energy and, hence, the electron is free to enter any of these two orbitals.

In the next 14 elements from cerium (Z=58) to lutecium (Z=71), the 4f orbital is successively filled pertaining to $[\text{Xe}]4f^15d^16s^2$ and $[\text{Xe}]4f^{14}5d^16s^2$ configurations, respectively. But you should remember, it is only Ce (Z=58), Gd (Z=64) and Lu (Z=71) that 5d orbitals have one electron while in all the remaining lanthanides the 5d orbitals remain vacant.

After lutecium, successive electrons occupy 5d orbitals and the electronic configuration builds up from $[\text{Xe}]4f^{14}5d^26s^2$ for hafnium to $[\text{Xe}]4f^{14}5d^{10}6s^2$ for mercury, the homologue of zinc and cadmium. Again, a minor departure from a steady increase in the number of d electrons occurs. For example, gold has $[\text{Xe}]4f^{14}5d^{10}6s^1$ configuration instead of $[\text{Xe}]4f^{14}5d^96s^2$. This, as you can see, has to do with the greater stability of half-filled/fully-filled orbitals. Finally the period is completed with successive occupation of the 6p orbitals from thallium, $[\text{Xe}]4f^{14}5d^{10}6s^26p^1$ to radon, $[\text{Xe}]4f^{14}5d^{10}6s^26p^6$.

Period 7: This period is still incomplete and contains 23 elements from francium (Z=87) to unnilennium (Z=109). In these elements, electrons are filled in 7s, 5f and 6d orbitals. Francium ($[\text{Rn}]7s^1$), radium ($[\text{Rn}]7s^2$) and actinium ($[\text{Rn}]6d^17s^2$) have electronic configurations analogous to those of caesium, barium and lanthanum, respectively. Thorium has the configuration $[\text{Rn}]6d^27s^2$. Thereafter, in next 13 elements from protactinium (Z=91) to lawrencium (Z=103) filling of 5f orbitals takes place successively. However, out of these, only Pa (Z=91), U (Z=92), Np (Z=93), Cm (Z=96) and Lr (Z=103) have an electron in 6d orbitals. In the rest of the elements, the 6d orbitals remain vacant. Thus, the electronic configuration of Lr (Z=103) is $[\text{Rn}]5f^{14}6d^17s^2$. The next six known elements of this period are members of 6d transition series which have the configurations $[\text{Rn}]5f^{14}6d^27s^2$ to $[\text{Rn}]5f^{14}6d^77s^2$.

Thus, we have examined the electronic configurations of elements in the periodic table. You can see from Table 1.4 that the elements occupying the same group of the periodic table have the same valence-shell electronic configuration. In other words, the elements having the same valence-shell electronic configuration recur periodically, that is after intervals of 2, 8, 8, 18, 18 and 32 in their atomic numbers. Therefore, periodicity in the properties of elements can be easily understood.

SAQ 4

a) What principles or rules are violated in the following electronic configurations? Write the name of the principle or rule in the space provided alongside each configuration.

- i) $1s^22s^3$
- ii) $1s^22s^22p_x^22p_y^1$
- iii) $1s^22p_x^2$

b) In the space provided write the electronic configuration of the atoms whose atomic numbers are

- i) 21
- ii) 24
- iii) 29

Electronic Configuration of Ions .

So far you have studied the electronic configuration of neutral atoms of elements. At this stage you will be naturally interested in knowing the electronic configuration of ions that are obtained by removal of electrons from the elements. When a gaseous iron atom having $[\text{Ar}]3d^64s^2$ ground state electronic configuration loses an electron, the Fe^+ ion is formed. This ion has its minimum energy in the configuration $[\text{Ar}]3d^7$, although the isoelectronic manganese atom has the configuration $[\text{Ar}]3d^54s^2$ in the ground state. Similarly, the ground states of the Fe^{2+} and Fe^{3+} ions are $[\text{Ar}]3d^6$ and $[\text{Ar}]3d^5$, respectively rather than $[\text{Ar}]3d^54s^1$ and $[\text{Ar}]3d^34s^2$ which are the ground states of isoelectronic atoms of chromium and vanadium, respectively. Evidently the differences in nuclear charge between Fe^+ and Mn , Fe^{2+} and Cr , and Fe^{3+} and V are important in determining the orbitals to be occupied by the electrons. However, along a series of ions carrying the same charge, the electronic configuration often changes much more regularly than the electronic configuration of the corresponding atoms. Thus, for dipositive ions from Sc^{2+} to Zn^{2+} , the ground state electronic configuration changes regularly from $[\text{Ar}]3d^1$ to $[\text{Ar}]3d^{10}$. For tripositive ions, there is a similar regular change from $[\text{Ar}]$ for Sc^{3+} to $[\text{Ar}]3d^9$ for Zn^{3+} . For tripositive ions of lanthanide elements, there is a regular change from $[\text{Xe}]4f^1$ for Ce^{3+} to $[\text{Xe}]4f^{14}$ for Lu^{3+} . Since the chemistry of elements is essentially that of their ions, the regularities in configuration of ions are much more important than the irregularities in the electronic configuration of the neutral atoms.

1.6 ELECTRONIC CONFIGURATION AND TYPES OF ELEMENTS

In the previous section you have studied the electronic configuration of the elements. Now we shall discuss the division of elements into various types depending upon their electronic configuration. Elements are thus, divided into four types, viz., *s*, *p*, *d* and *f* block elements depending upon the nature of the atomic orbitals into which the differentiating or the last electron enters.

s-Block elements: In these elements, the differentiating electron enters the '*ns*' orbital. Alkali and alkaline earth metals of Groups 1 (IA) and 2 (IIA) belong to this block. You have already seen that their valence shell electronic configurations are ns^1 and ns^2 , respectively. You have also studied that each period of the periodic table begins with an alkali metal. All the elements of this block are metals.

p-Block elements: In the elements belonging to this block, the *p*-orbitals are successively filled. Thus, the elements of the Groups 13 (IIIA), 14 (IVA), 15 (VA), 16 (VIA), 17 (VIIA) and 18 (Zero) are members of this block, since in the atoms of these elements, the differentiating electron enters the *np* orbitals. The *ns* orbitals in the atoms of these elements are already completely filled so they have the valence shell electronic configuration ns^2np^{1-6} .

The elements of *s*- and *p*-blocks are also known as normal, representative or main group elements.

d-Block elements: The elements in which the differentiating electron enters the $(n-1)d$ orbitals are called *d*-block elements. These elements are placed in the middle of the periodic table between the *s*- and *p*- block elements. You can represent the configuration of the atoms of the elements of this block by $(n-1)d^{1-10}ns^{0,1,2}$. These elements which are also called transition elements are divided into four series corresponding to the filling of *3d*-, *4d*-, *5d*-, or *6d*- orbitals. While the *3d*, *4d* and *5d* series consist of 10 elements each, the *6d* series is incomplete and has only seven elements, viz., Ac ($Z=89$) and from Unq ($Z=104$) to Une ($Z=109$). The elements from Sc ($Z=21$) to Zn ($Z=30$), Y ($Z=39$) to Cd ($Z=48$), La ($Z=57$) and from Hf ($Z=72$) to Hg ($Z=80$) are the members of *3d*, *4d*, and *5d* series, respectively.

d-Block elements are also known as transition elements.

f-Block elements: The elements in which the extra electron enters $(n-2)f$ orbitals are called the *f*-block elements. The atoms of these elements have the general configuration $(n-2)f^{1-14}(n-1)d^{0,1}ns^2$. These elements belong to two series depending upon the filling of *4f* and *5f* orbitals. Elements from Ce ($Z=58$) to Lu ($Z=71$) are the members of the *4f* series while those from Th ($Z=90$) to Lr ($Z=103$) belong to the *5f* series. Elements of *4f* series

which follow lanthanum in the periodic table are known as **lanthanides** whereas those of 5f series following actinium are called **actinides**. All these elements are collectively referred to as **inner-transition** elements because of filling of electrons in an inner (n-2)f subshell.

f-Block elements are also known as inner-transition elements.

SAQ5

- a) Name the blocks into which the elements in the periodic table have been divided.
-
-
-
- b) Differentiate between transition and inner-transition elements in terms of their electronic configuration.
-
-
-
-

1.7 SUMMARY

In this unit you studied the development of the periodic table and the periodic law. We have followed the development of modern periodic law from the earliest classification of elements into metals and non-metals, to Dobereiner's Triads, Newlands' Law of Octaves, Chancourtois' Telluric Screw, Lothar Meyer's curves and Mendeleev's Periodic Law.

The periodic table organises the elements into periods and groups in order of increasing atomic number so as to display the periodic law. The periodic law can be stated as "the properties of elements are the periodic function of their atomic number". This periodicity in properties arises due to the periodicity in the electronic configuration of the elements. Thus, the elements occurring in the same group of periodic table have similar electronic configurations and depict similar physical and chemical properties.

Electronic configurations of atoms can be determined with the help of *aufbau* principle, Pauli's exclusion principle and Hund's rule of maximum multiplicity.

On the basis of their electronic configurations, elements are divided into four categories. Those having $ns^{1,2}$ configuration are called s-block elements. Elements of Groups 1 and 2 are the members of this block. Elements having ns^2np^{1-6} configuration are known as p-block elements. Elements of Groups 13 to 18 belong to this class. d-Block elements have $(n-1)d^{1-10}ns^{0,1,2}$ configuration and thus, elements of Groups 3 to 12 belong to this block. f-Block elements have $(n-2)f^{1-14}(n-1)d^{0,1}ns^2$ configuration. Lanthanide and actinide elements are members of this block.

1.8 TERMINAL QUESTIONS

1. Newlands attempted to show that there was a periodic similarity in the properties of every eighth element. Yet today, we see that for the elements of period 2 and 3, it occurs at every ninth element. Explain.
2. What are the advantages of classifying elements into groups?
3. State the modern periodic law. What is the cause of periodicity in the properties of the elements?
4. Explain how elements are arranged in the long form of the periodic table.
5. Explain what is meant by a group and a period.
6. Explain why there are only two elements in the first period, eight elements each in the second and third periods, eighteen elements each in the fourth and fifth periods, thirty-two elements in the sixth period of the periodic table.
7. Explain *aufbau* principle and *n+l* rule.

8. Explain Pauli's exclusion principle.
9. State Hund's rule of maximum multiplicity.
10. Explain how the elements are classified into *s*, *p*, *d* and *f*-blocks.

1.9 ANSWERS

Self Assessment Questions

- 1 a) Mendeleev used atomic weight to classify elements. He arranged the elements in order of increasing atomic weight in his periodic table.
- b) Some of the defects of Mendeleev's periodic table are:
 - i) Anomalous position of hydrogen.
 - ii) No place for isotopes of elements.
 - iii) No proper place for lanthanides and actinides.
 - iv) In some cases, elements of higher atomic weights were placed before those having lower atomic weights, e.g., Co before Ni and Te before I.
 - v) He could not predict the existence of noble gases.

Property	K	Sc	Prediction for Ca	Observed for Ca
At. weight	39.1	41.6	40.3	40
Density (10^3 kg m^{-3})	0.86	3.0	~1.9	1.55
Melting Point (K)	336.7	1812	1074	1111
Formula of oxide	K_2O	Sc_2O_3	CaO	CaO
Formula of chloride	KCl	ScCl_3	CaCl_2	CaCl_2
Nature of chloride	Ionic	Ionic	Ionic	Ionic
Reactivity towards water and acids	Reacts vigorously evolving H_2	Reacts when finely divided or heated evolving H_2	Should be moderately reactive and should liberate H_2	As predicted

- 2 a) Group 18 containing noble gases appears in the modern periodic table, but it did not appear in Mendeleev's original table because noble gases were not discovered at that time.

b) In periodic table, elements are arranged in the order of increasing atomic number.

Atomic Number	IUPAC name	Symbol
107	Unnilseptium	Uns
108	Unniloctium	Uno
109	Unnilennium	Une
110	Ununnilium	Uun
111	Unununium	Uuu
112	Ununbium	Unb
113	Ununtrium	Unt
114	Ununquadium	Unq

- 4 a) i) Exclusion principle ii) Hund's rule iii) *aufbau* principle

b) i) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$ or $[\text{Ar}] 3d^1 4s^2$

ii) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$ or $[\text{Ar}] 3d^5 4s^1$

iii) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ or $[\text{Ar}] 3d^{10} 4s^1$

iv) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$ or $[\text{Ar}] 3d^{10} 4s^2 4p^5$

v) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 5d^1 6s^2$ or $[\text{Xe}] 5d^1 6s^2$

- 5 a) In the periodic table, elements are divided into four blocks, namely, *s*-, *p*-, *d*- and *f*-block elements. Elements of *s*- and *p*-blocks are representative elements whereas those of *d*- and *f*-blocks are known as transition and inner-transition elements.

b) Transition elements are called *d*-block elements. In these elements, the differentiating electron is added to the *d*-orbitals of the penultimate shell. Their

electronic configuration can be represented by $(n-1)d^{1-10}ns^{0-2}$. Inner-transition elements are known as *f*-block elements. In these elements the differentiating electron enters the *f*-orbitals of the shell which is second from the outermost shell. They have the general electronic configuration $(n-2)f^{1-14}(n-1)d^{0,1}ns^2$.

Terminal Questions

- 1 In Newlands' time, the noble gas elements were not known. He, therefore, reported a periodic similarity in properties of every eighth element. Obviously with the inclusion of noble gases in the periodic table, for the elements of period 2 and 3 this similarity will occur at every ninth element.
- 2 The necessity of classifying elements into groups was to facilitate the study of the vast chemistry of a large number of elements and their compounds.
- 3 Modern periodic law can be stated as "the properties of the elements are the periodic function of their atomic numbers". Cause of periodicity in the properties of elements lies in the electronic configuration of the elements. Elements having similar valence-shell electronic configuration recur at regular intervals in their atomic number, therefore, they exhibit similarities in properties.
- 4 In the long form of the periodic table elements are arranged in seven horizontal rows and eighteen vertical columns in the order of increasing atomic number in such a manner that elements having similar properties occur in the same vertical column.
- 5 Vertical columns and horizontal rows in the periodic table are known as groups and periods, respectively. Elements of a particular group in the table have similar properties and differ from the elements of other groups in the table. Properties of elements in a period vary gradually from one element to another.
- 6 Answer to this lies in the electronic configuration of the atoms of the elements. Each period of the table begins with an element in the atoms of which filling of *s*-orbitals of a new principal shell starts. Except for the first period which ends with the completion of 1*s* orbital in noble gas helium, all others end with noble gas elements in whose atoms the *p*-orbitals of the same principal shell get completely filled. Thus, the first period contains only two elements, hydrogen and helium pertaining to 1*s*¹ and 1*s*² configuration. The second period contains eight elements in the atoms of which 2*s* and 2*p* orbitals are filled. The third period also contains eight elements due to filling of 3*s* and 3*p* orbitals. The fourth period begins with filling of 4*s* orbital and ends with the completion of 4*p* orbitals. But in between 4*s* and 4*p* orbitals, 3*d* orbitals are also filled. Thus, the fourth period contains eighteen elements. Similarly, in the elements of fifth period the 5*s*, 4*d*, and 5*p* orbitals are filled, so it also contains eighteen elements. The sixth period contains thirty-two elements pertaining to the filling of 6*s*, 4*f*, 5*d* and 6*p* orbitals. Period 7 can also accommodate an equal number of elements, but it is incomplete and contains only twenty-three elements pertaining to the complete filling of 7*s* and 5*f* orbitals and partial (7) filling of 6*d* orbitals.
- 7 Electronic configuration of atoms can be derived with the help of *aufbau* principle. According to this principle, electrons are filled in orbitals one at a time in order of increasing energy. Order of filling of orbitals is 1*s* → 2*s* → 2*p* → 3*s* → 3*p* → 4*s* → 3*d* → 4*p* → 5*s* → 4*d* → 5*p* → 6*s* → 4*f* → 5*d* → 6*p* → 7*s* → 5*f* → 6*d*. Order of filling of electrons in orbitals can be remembered in the form of *n+l* rule according to which the subshell with the lowest value of *n+l* fills first. When two subshells have the same value of *n+l*, the subshell with lower value of *n* fills first.
- 8 Pauli's exclusion principle states that no two electrons in an atom can have the same values of all the four quantum numbers. In other words, an orbital can accommodate at the most two electrons.
- 9 Hund's rule of maximum multiplicity states that as far as possible in a given atom in the ground state, electrons in the same subshell will occupy different orbitals and will have the parallel spins.
- 10 Depending upon the filling of electrons in *s*, *p*, *d* or *f* orbitals, elements are classified into *s*, *p*, *d* or *f*-blocks in the periodic table.