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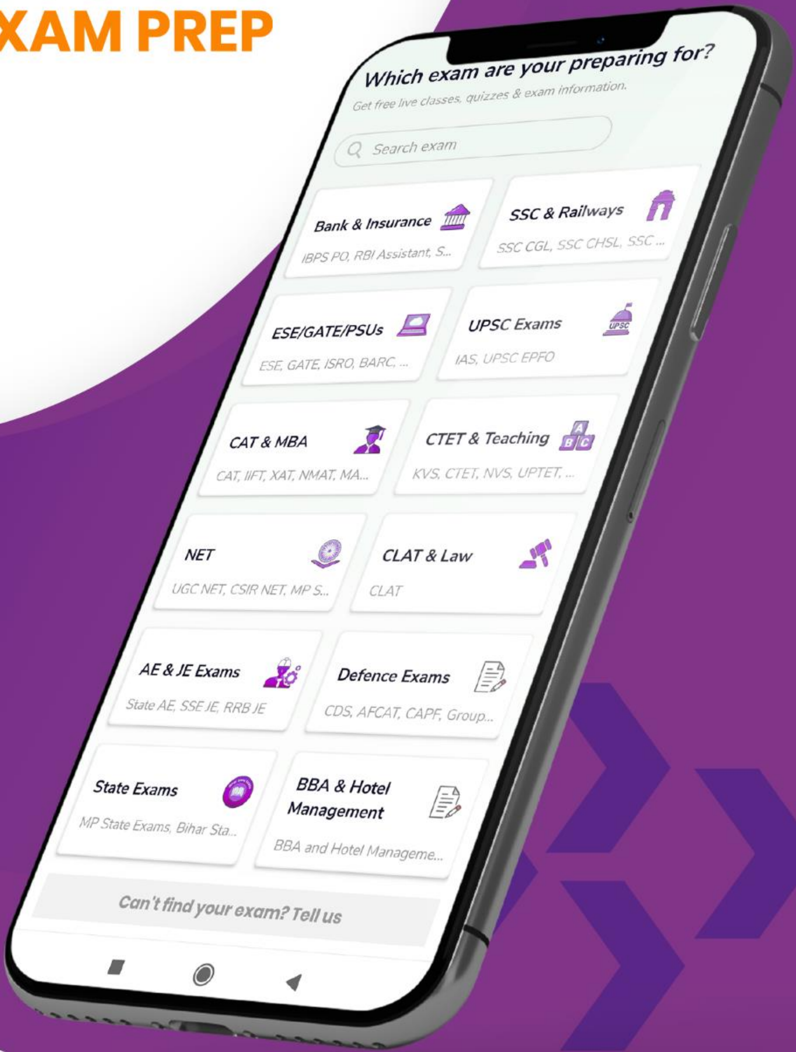

Chemistry Notes on Periodic Table

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The periodic table of elements is frequently used in the area of chemistry to look up chemical elements because it is organised in such a way that it shows periodic trends in the chemical characteristics of the elements. However, the Periodic table usually simply shows the element's symbol and not the complete name.

Here is all you need to know about the Periodic Table. As per the latest exam trend, periodic table questions carry a good weightage of marks. To help you update your knowledge, we have highlighted the essential pointers of the Periodic Table.

Introduction

We have 118 elements in the Periodic Table. Studying each element, its isolation, and its physical and chemical properties would be an arduous task. A systematic arrangement of the element according to their electronic configuration and thereafter study of their properties both physical and chemical forms the hallmark of this chapter. Elements are classified according to groups and periods for a better understanding of the properties of the elements. Look at the table below for a complete understanding of Atomic No., their mass along with Element Name & Symbol.

Atomic No.	Atomic Mass	Element Name	Symbol
1	1.0079	Hydrogen	H
2	4.0026	Helium	He
3	6.941	Lithium	Li
4	9.0122	Beryllium	Be
5	10.811	Boron	B
6	12.0107	Carbon	C
7	14.0067	Nitrogen	N
8	15.9994	Oxygen	O
9	18.9984	Fluorine	F
10	20.1797	Neon	Ne
11	22.9897	Sodium	Na
12	24.305	Magnesium	Mg
13	26.9815	Aluminium	Al
14	28.0855	Silicon	Si
15	30.9738	Phosphorus	P
16	32.065	Sulphur	S
17	35.453	Chlorine	Cl
18	39.948	Argon	Ar
19	39.0983	Potassium	K
20	40.078	Calcium	Ca
21	44.9559	Scandium	Sc
22	47.867	Titanium	Ti
23	50.9415	Vanadium	V
24	51.9961	Chromium	Cr
25	54.938	Manganese	Mn
26	55.845	Iron	Fe
27	58.9332	Cobalt	Co

28	58.6934	Nickel	Ni
29	63.546	Copper	Cu
30	65.39	Zinc	Zn
31	69.723	Gallium	Ga
32	72.64	Germanium	Ge
33	74.9216	Arsenic	As
34	78.96	Selenium	Se
35	79.904	Bromine	Br
36	83.8	Krypton	Kr
37	85.4678	Rubidium	Rb
38	87.62	Strontium	Sr
39	88.9059	Yttrium	Y
40	91.224	Zirconium	Zr
41	92.9064	Niobium	Nb
42	95.94	Molybdenum	Mo
43	98	Technetium	Tc
44	101.07	Ruthenium	Ru
45	102.9055	Rhodium	Rh
46	106.42	Palladium	Pd
47	107.8682	Silver	Ag
48	112.411	Cadmium	Cd
49	114.818	Indium	In
50	118.71	Tin	Sn
51	121.76	Antimony	Sb
52	127.6	Tellurium	Te
53	126.9045	Iodine	I
54	131.293	Xenon	Xe
55	132.9055	Cesium	Cs
56	137.327	Barium	Ba
57	138.9055	Lanthanum	La
58	140.116	Cerium	Ce
59	140.9077	Praseodymium	Pr
60	144.24	Neodymium	Nd
61	145	Promethium	Pm
62	150.36	Samarium	Sm
63	151.964	Europium	Eu
64	157.25	Gadolinium	Gd
65	158.9253	Terbium	Tb
66	162.5	Dysprosium	Dy
67	164.9303	Holmium	Ho
68	167.259	Erbium	Er
69	168.9342	Thulium	Tm
70	173.04	Ytterbium	Yb
71	174.967	Lutetium	Lu
72	178.49	Hafnium	Hf
73	180.9479	Tantalum	Ta
74	183.84	Tungsten	W
75	186.207	Rhenium	Re

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76	190.23	Osmium	Os
77	196.9665	Iridium	Ir
78	192.217	Platinum	Pt
79	195.078	Gold	Au
80	200.59	Mercury	Hg
81	204.3833	Thallium	Tl
82	207.2	Lead	Pb
83	208.9804	Bismuth	Bi
84	209	Polonium	Po
85	210	Astatine	At
86	222	Radon	Rn
87	223	Francium	Fr
88	226	Radium	Ra
89	227	Actinium	Ac
90	232.0381	Thorium	Th
91	231.0359	Protactinium	Pa
92	238.0289	Uranium	U
93	237	Neptunium	Np
94	244	Plutonium	Pu
95	243	Americium	Am
96	247	Curium	Cm
97	247	Berkelium	Bk
98	251	Californium	Cf
99	252	Einsteinium	Es
100	257	Fermium	Fm
101	258	Mendelevium	Md
102	259	Nobelium	No
103	262	Lawrencium	Lr
104	261	Rutherfordium	Rf
105	262	Dubnium	Db
106	266	Seaborgium	Sg
107	264	Bohrium	Bh
108	277	Hassium	Hs
109	268	Meitnerium	Mt
110	261.9	Darmstadtium	Ds
111	271.8	Roentgenium	Rg
112	285	Copernicium	Cn
113	286	Nihonium	Nh
114	289	Flerovium	Fl
115	290.196	Moscovium	Mc
116	293	Livermorium	Lv
117	294	Tennessee	Ts
118	294	Oganesson	Og

Significance and Brief History of the Development of Periodic Table

The elements are the basic units of all types of matter. Only 31 elements were known in 1800. Now more than 115 elements are known. It is impossible to remember the properties of each element and its compounds. Therefore, many attempts have been made to classify elements into fewer groups. The purpose of the classification has been to make the study of the chemistry of elements and their compounds easier. Dmitri I. Mendeleev (1834-1907) developed the periodic table. From the relationships embodied in the table, he predicted the existence as well as properties of elements then unknown. These predictions came out to be amazingly accurate. During the course of time, the basis of classification changed from atomic mass to atomic number.
The study of chemistry has become simple on the basis of the modern classification of elements. Now, by knowing the properties of one element of a group, it is possible to predict the properties of the other members. One need not remember all the properties of elements or their compounds. All one has to do is to know the trends of properties in a group and in a period of the periodic table.

Brief History of the Development of Periodic Table

In 1829, Döbereiner suggested that elements could be arranged in groups of three i.e. triads, in which the atomic weight of the middle element was nearly the mean of the atomic weights of the other two. However, only a limited number of electrons could be grouped into these triads.

Dobereiner's Triads of Elements			
	Triads		**Mean atomic weight**
Lithium (7)	Sodium(23)	Potassium(39)	(7+ 39)/2 = 23
(Li)	(Na)	(K)	
Calcium (40)	Strontium (87.5)	Barium (137.5)	(40 +137.5)/2 = 88.75
(Ca)	(Sr)	(Ba)	
Phosphorus (31)	Arsenic (76)	Antimony (120)	(31 + 120)/2 = 75.5
(P)	(As)	(Sb)	
Sulphur (32)	Selenium (79)	Tellurium (127.5)	(32+ 137.5)/2 = 79.25
(S)	(Se)	(Te)	
Chlorine (35.5)	Bromine (80)	Iodine (127)	(35.5 + 127)/2 = 81.25
(Cl)	(Br)	(I)	

John A.R. Newlands, in 1865-1866, reported that if the elements were arranged in order of their increasing atomic weights, the eighth element starting from a given one, possessed properties similar to the first, like the eighth note in an octave of music. He called it the law of octaves. It worked well for the lighter elements but failed when applied to heavier elements. In 1869, J. Lothar Meyer in Germany and Dmitri I. Mendeleev in Russia, working independently, gave a more detailed and accurate relationship among the elements. Lother plotted atomic volumes (=atomic mass/volume) versus atomic weights of elements and obtained a curve. He pointed out that, elements occupying similar positions in the curve possessed similar properties.

Mendeleev's Periodic Table

In March 1869, Mendeleev gave his famous scheme of the periodic classification of elements. "It states that the properties of the elements are the periodic functions of their atomic weights.

Mendeleev arranged all the elements known at that time in horizontal rows in order of increasing atomic weights. He left some gaps in the table for undiscovered elements and also

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predicted their properties. Eventually, when these elements were discovered, it was amazing to find that they fit correctly in the table. For example, as can be seen in Mendeleev's original periodic table there are two gaps between zinc (group II) and arsenic (group V). As these undiscovered elements were to follow aluminium and silicon, Mendeleev named them eka-aluminium and eka-silicon. When these elements, now known as gallium (eka-aluminium) and germanium (eka-silicon), were discovered, their properties were agreed very well with those given by Mendeleev.

Mendeleev's prediction for eka-aluminium (Gallium)

Property	**Eka-aluminium (Ea) predicted in 1871 by Mendeleev**	**Gallium (Ga)**	
		Reported in 1875	**Currently accepted**
Atomic mass/amu	68	69.9	69.72
Density, r/ /g cm^{-2}	5.9	5.94	5.904
Melting point, T/K	Low	303.15	302.78
Solubility in acid and alkali	Ea will dissolve slowly in both acid and alkali	Ga dissolves slowly in both acid and alkali	Ga dissolves slowly in both acid and alkali
Formula of oxide	Ea_2O_3	Ga_2O_3	Ga_2O_3 ((a and b forms)
Density of oxide r /g cm^{-2}	5.5	-	6.44 ((a form) and 5.883 (b form)
Reactions of sulphate	$Ea_2(SO_4)_3$ will form alum	Gallium forms alums	Gallium forms alums, e.g. $(NH_4)_2SO.Ga(SO).24HO$
Preparation of Sulphide	Ea_2S_3 will be precipitated by H_2S or $(NH_4)_2S$	Ga_2S_3 is precipitated by H_2S or $(NH_4)_2S$	Ga_2S_3 is precipitated by H_2S or $(NH_4)_2S$
Properties of chloride	$EaCl_3$ will be more volatile than $ZnCl_2$	$GaCl_3$ is more volatile than $ZnCl_2$	$GaCl_3$ is more volatile than $ZnCl_2$

Mendeleev's prediction for eka-silicon (Germanium)

Property	**Eka-silicon (Es) predicted in 1871 by Mendeleev**	**Germanium(Ge)**	
		Reported in 1886 by Bioshaudran	**currently accepted**
Atomic mass/ amu	72	72.32	72.59
Density, /g cm^{-3}	5.5	5.47	5.35
Melting point, T/K	High	-	1220
Specific heat capacity, c/J g^{-1}	0.3051	0.3177	0.3903
Molar Volume, V_m/ cm^3	13	13.22	13.5
Colour	Dark grey	Greyish white	Greyish white

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Valence	4	4	4
Reaction with acids and alkalis	Es will be slightly attacked by acids but will resist attack by alkalis	Ge is dissolved by neither HCl nor NaOH, but dissolved by conc. NaOH	Ge is dissolved by neither HCl nor NaOH, but dissolved by conc. NaOH
Boiling point of tetraethyl derivative, T/K	433	433	458-460
Density of the dioxide, /g cm^{-3}	4.7	4.7	4.2
Density of the tetrachloride, /g cm^{-3}	1.9	1.887	1.844
Boiling point of the tetrachloride, T/K	373	359	357

Modern Periodic Law and Present Form of Periodic Table

According to the modern periodic law, the properties of the elements and their compounds are a periodic function of their atomic numbers. Thus, in the modern periodic table, atomic number (which is equal to the nuclear charge) forms the basis of the classification of elements.

Present Form of Periodic Table

There are several forms of the periodic table. The most popular version is the long form. In that table, there are seven horizontal rows called periods. Each period starts with a new principal quantum number, n, and the electrons are filled up in orbitals according to the Aufbau principle. The first period has two elements, hydrogen ($1s^1$), and helium ($1s^2$), and the first shell (K) is completed. The second period starts with n = 2 and has eight elements. Starting with lithium ($2s^1$), it ends with neon ($2s^2\ 2p^6$) and thus completes the second shell (L). In the third period, shell M starts getting filled, (n=3), and also contains eight elements. It starts with sodium ($3s^1$) and completes at argon ($3s^2\ 3p^6$).

Long form of the periodic table of the elements with their atomic numbers and ground state electronic configurations (according to the latest 1984 IUPAC recommendation)

In contrast to the second period, the outer shell is not completely filled in with elements of the third period, and the orbitals of the 3d subshell remain vacant. Elements in which the s subshell is filled are called s-block elements and those in which the p subshell is filled are known as p-block elements.

In the fourth period, the N shell (n = 4) starts filling with potassium ($4s^1$) and is completed with krypton ($4s^2 4p^6$). It has 18 elements, i.e. ten more than in the third period. This is because of the elements in which the filling up of electrons in the 3d orbitals takes place after 4s orbital but before 4p orbitals:

$_{21}$Sc $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$

$_{22}$Ti $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$

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$_{30}Zn$ $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^{10}$ $4s^2$

Elements whose d orbitals are filled are called d-block elements. In d-block elements of the fourth period, the M shell is filled until it contains 18 electrons. Similarly, the fifth period (n = 5) has 18 elements. It starts with rubidium ($5s^1$) and ends with xenon ($5s^2$ $5p^6$). The sixth period (n = 6) contains 32 elements in which the electrons enter in 6s, 4f, 5d and 6p orbitals, in that order. It begins with two s elements (Cs and Ba) followed by lanthanum (La) in which the d orbitals of the penultimate shell begin to be filled. But immediately after this, we have 14 elements ($_{58}Ce$ to $_{71}Lu$) in which 4f orbitals are in the process of getting filled. Then, the 5d orbitals are filled (from $_{72}Hf$ to $_{80}Hg$) and finally, the period ends with six p block elements ($_{81}Tl$ to $_{86}Rn$). The seventh period (n = 7) contains two s block elements ($_{87}Fr$ and $_{88}Ra$), followed by a d block elements actinium ($_{89}Ac$), and 14f block elements ($_{90}Th$ to $_{103}Lr$) and again ends with d block elements (Z = 104 to 107). This is an incomplete period.

The vertical columns in the periodic table are called groups or families of elements. According to the latest IUPAC (International Union of Pure and Applied Chemistry) recommendation, the groups are numbered from 1 to 18.

Types of Elements

The elements can be classified into four types depending on their electronic configurations (Table 5.4). Thus, we have:
1. Noble gases
2. Representative elements (s-and p-block elements)
3. Transition elements (d-block elements)
4. Inner transition elements (f-block elements)

Noble Gases

Noble gases constitute group 18 of the periodic table. Except for the first element of the group, helium (which has $1s^2$ configuration) all other elements, namely neon, argon, krypton, xenon and radon have ns^2 np^6 electronic configuration in the outermost shell. Because of the stable arrangement of electrons in these elements, they exhibit a very low chemical reactivity.

Representative Elements

Elements belonging to group 1 (alkali metals; Li ... Fr) with outermost electronic configuration ns^1 and those belonging to group 2 (alkaline earth metals; Be... Ra) with outermost electronic configuration ns^2 are placed in the s block. Elements of groups 13 to 17 (outermost electronic configuration varying from ns^2 np^1 to ns^2 np^5) belong to the p block of the periodic table. The elements of s and p blocks are collectively called representative elements.

Noble gases are at the end of each period of the representative elements; these are also grouped with representative elements. The chemistry of these elements depends on the number of electrons in the valence shell (outermost shell). For groups 1 and 2, the group number indicates the number of valence electrons. Subtracting 10 gives the number of valence electrons for groups 13 to 17 from the group number.

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Transition Elements

Elements belonging to groups 3 to 12 (in the middle of the periodic table) with an outer electronic configuration (n - 1) d^{1-10} ns^{1-2} constitute the d block of the periodic table and are called transition elements. In these elements, n is 4, 5 or 6 with the corresponding filling of 3d, 4d or 5d orbitals. They are all metals and are characterized by variable oxidation states, the formation of coloured ions and complexes.

Inner Transition Elements

At the bottom of the periodic table, there are two rows-one of lanthanides or lanthanide series (z = 58 to 71) and the other of actinides or actinide series (Z = 90 to 107) containing incomplete 4f and 5f orbitals respectively. They also have incomplete (n - 1) d orbitals and are characterized by the outer electronic configuration (n - 2) f^{1-14} (n - 1) d^{0-1} ns^2. As in each of these series, an inner f electron is added to each element, the two series of elements are called f-block elements or inner transition elements. Among themselves, lanthanides and actinides show similar properties. All the elements of the two series are metals.

Through this periodic classification, we can organize and systematize the study of elements and their compounds. It is seen that the electronic configurations of elements are directly or indirectly related to their physical as well as chemical properties. From this classification, we can understand the cause of periodicity in properties and general trends in the behaviour of elements.

Periodic Trends in Properties of Elements

According to the modern periodic law, properties of elements are repeated at certain intervals or periods when the elements are arranged in the order of increasing atomic number. In a period, the number of valence electrons increases with an increase in atomic number. There is a repetition of similar electronic configurations in the outermost shells in a group. Therefore, in a group, properties are similar. In the long form of the periodic table, elements with the same number of outer electrons are in the same group. Moreover, there is a gradation in the properties of elements in a group. For example, all members of group 1 have one electron in the outermost shell. Properties of these elements are similar to each other. Similarly, F, Cl, Br and I have seven electrons in their outermost shells. They are chemically alike. Members of group 18 have similar electronic configurations with 8 electrons in their outermost shell (except He). All noble gases have similar properties. We shall now discuss periodic trends in some of the properties of elements.

Density, Melting Point and Boiling Point

The properties like density, melting point and boiling point also show gradation, though not very regular, in a group of the periodic table. In a period of representative elements, density increases across a period and reaches a maximum value somewhere in the middle. Melting points and boiling points increase, attain maximum values in the middle, thereafter, and begin to decline.

Ionization Energy

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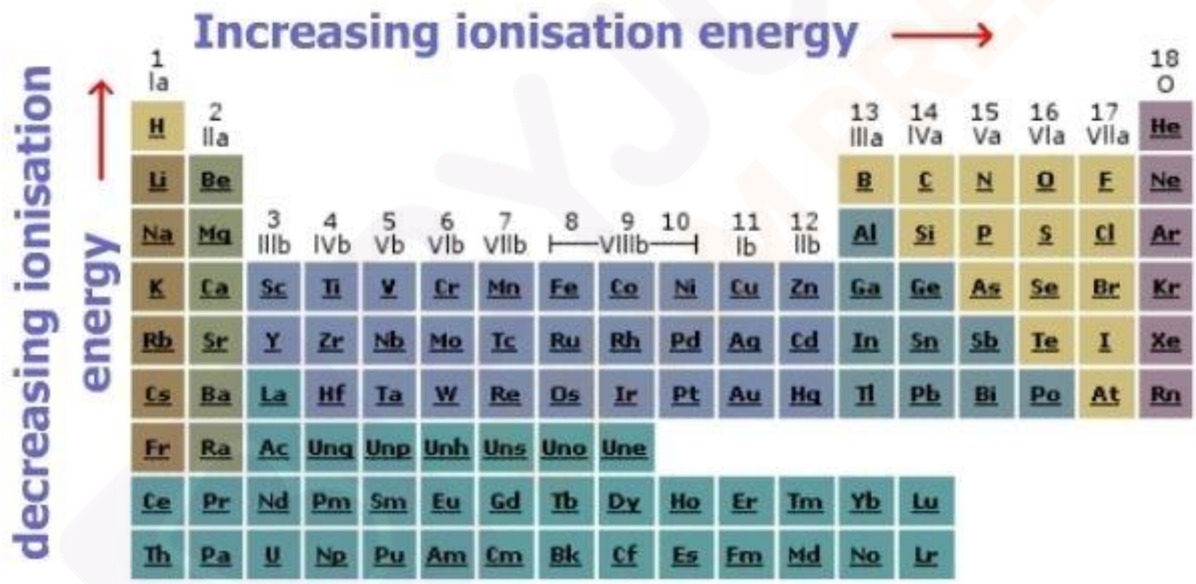
First ionization energy is the energy required to remove an electron from an isolated, neutral gaseous atom of an element resulting in the formation of a positive ion.
$M(g)$ + Energy ------> $M^+(g) + e^-$
For example, $Li(g)$ + 520 kJ mol^{-1} ------> $Li^+(g) + e^-$

Thus, the first ionization energy of lithium is 520 kJ mol^{-1}. If another electron is to be removed, much more energy (called second ionization energy) will be needed.
Li^+ + 7297 kJ mol^{-1} -------> $Li^{2+} + e^-$

This is because of the fact that it is more difficult to remove an electron if the atom already possesses a positive charge. Thus, the successive values of the ionization energy of an element will show an increasing trend.

The below figure shows a plot of first ionization energies of some elements against their respective atomic numbers. Certain trends are quite evident from the plot:
1. Noble gases occupy peaks because of their extremely stable configurations.
2. Alkali metals possess the lowest values of ionization energy showing thereby that they are highly reactive.



3. In general, the values increase across a period (e.g. Li to Ne). This is due to the increasing nuclear charge and the attractive forces of the nucleus which strongly hold the electrons.
4. Within a group (e.g. Li to Cs), ionization energies become progressively smaller as we move down the column. This is because of the fact that the outer electron, which is being removed lies farther from the nucleus and thus it becomes easier to pull it out.

Exceptions

1. **Ionisation Be > Ionisation B** : The ionisation energy of Beryllium is greater than the ionisation energy of Boron due to penetration effect. **Be= 1s2 2s2** & **B=1s2 2s2 2p1** The S orbital is more closer to nucleus and hence require more energy !
2. **Ionisation N > Ionisation O : O= 1s2 2s2 2p4 & N=1s2 2s2 2p3** The Ionisation energy of Nitrogen is greater than the ionisation energy of oxygen due to completely half-filled and the stable electron configuration.

General trends in ionization energies in relation to the periodic table may be depicted as shown in the figure. Ionization energy is either determined by means of a discharge tube method or by the spectroscopic method

Electron Affinity

Electron affinity is the energy released when an electron is added to an isolated atom in the gaseous state. For example,

$X(g) + e^- \text{------>} X^-(g) + \text{Energy}$
(where X = F, energy 328 kJ/mol; X = Cl, energy = 349 kJ/mol; X = Br, energy = 324 kJ/mol; X = I, energy = 295 kJ/mol).

Halogens (elements of group 17) can take up an electron to acquire the stable noble gas configuration. Their values for electron affinity are thus very high. As we move across a period, electron affinity usually increases, and while going down a group it decreases. The electron affinity of an element is generally determined indirectly from thermodynamic data.
Although electron affinities of all the elements have not been determined, the following trends in electron affinities of some elements in the periodic table are evident.

Electron affinities generally decrease in moving down the group. This is expected on account of the increase in the size of the atom on moving down the group. Due to an increase in the size of atoms, the effective nuclear attraction for electrons decreases. As a result, there are fewer tendencies to attract additional electrons with an increase in atomic number.
In the figure electron affinities of halogens are plotted against their atomic numbers. It may be noted that contrary to expectation, the electron affinity of fluorine is lower than that of chlorine. This is because the fluorine atom has a very compact electronic shell due to its small size. The compactness of the fluorine shell results in electron-electron repulsion whenever an electron is introduced into its 2p shell. This is why its electron affinity is less than expected. In the chlorine atom, the 3p orbitals are not as compact as the 2p orbitals in fluorine atom. The chlorine atom, because of weaker electron-electron repulsion, more readily accepts the incoming electron. The electron affinity of chlorine is, therefore, higher than that of fluorine.

In the case of noble gases, the outer s and p orbitals are completely filled. No more electrons can be accommodated in these orbitals. Noble gases, therefore, show no tendency to accept electrons. Their electron affinities are zero.
Electron affinities generally increase as we move across a period from left to right. This is due to the increase in the nuclear charge, which results in greater attraction for electrons. In the second period, for example, the electron affinity has a maximum value for fluorine. The second electron affinity refers to a process in which an electron is added to a negative ion. For example

$O^-(g) + e^- \text{----->} O^{2-}(g)$

Since a negative ion (O^-) and electron repel each other, energy is required and not released by the process. All second affinities are thus negative.

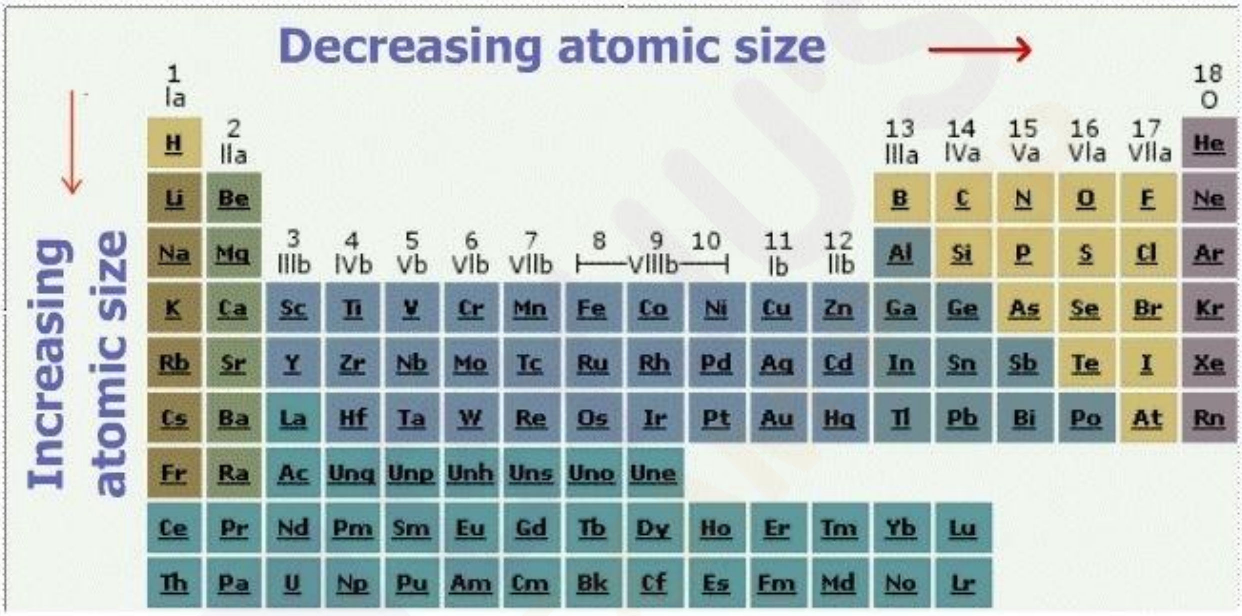
Electronegativity

Electronegativity is the tendency of an atom to attract the shared pair of electrons in a chemical bond towards itself. It is, thus, a measure of the ability of an atom in a molecule to attract electrons. In general, electronegativity increases from left to right across any period and from bottom to top in any group of the periodic table. The most highly electronegative elements

(Fluorine) are found in the upper right corner of the periodic table (ignoring the noble gases.) The least electronegative elements (Caesium) are found in the lower-left corner of the table.

When two elements of widely different electronegativities combine, an ionic compound results. The electronegativity values of the non-metals do not differ much. The bonds formed between non-metals are essentially covalent with some polar character. The electronegativity difference gives an indication of the degree of polarity of the covalent bonds. If the difference is zero or very small, an essentially non-polar bond with equal or almost equal sharing of electrons can be assumed. The larger the electronegativity difference, the more polar the covalent bond is. The bond is polarized in the direction of the atom with larger electronegativity.

Atomic radii: The distance between the nucleus and the outermost shell containing the electron cloud is known as the atomic radius.



In general, the atomic radius decreases while going across a period in the periodic table Figure. This is because of the increasing nuclear charge while the electrons are added to the same shell. On the other hand, the atomic radius increases as we move from top to bottom in a group of the periodic table Figure. This happens because of the increasing number of shells. Though the nuclear charge also increases, the effect of adding a new shell is very large and overcomes the effect of increased nuclear charge.
Variation of the atomic radius with an atomic number for alkali metals (group 1) and halogens (group 17).

Ionic radii: It is defined as the distance between the nucleus and the point up to which the element has influence over its electron cloud. Generally, the cationic radius is smaller than its neutral atom, whereas the anionic radius is larger. An important point to remember here is ions which are having same no. of electrons are called isoelectronic species (eg. S^{2-} K^{+} $Ca2^{+}Cl^{-}$).

Valency

From the position of an element in a particular group of the periodic table, one can infer its valency. The valency of a representative element is usually given by the number of electrons in the outermost orbital and /or equal to eight minus the number of outermost electrons. For example, alkali metals (group 1) having one outermost electron are monovalent and alkali earth metals (group 2) having two outermost electrons are bivalent. Halogens (group 17) with seven outermost electrons are monovalent (eight minus number of outermost electrons). There are, however, some exceptions to this rule.

Group	1	2	13	14	15	16	17	18
Number of valence electron	1	2	3	4	5	6	7	8
Valence	1	2	3	4	3,5	2,6	1,7	0,8

Periodic Trends in Valence of Elements as shown by the Formulas of Their Compounds.

Group	1	2	13	14	15	16	17
Formula of hydride	LiH NaH KH	CaH_2	B_2H_6 AlH_3	CH_4 SiH_4 GeH_4 SnH_4	NH_3 PH_3 AsH_3 SbH_3	H_2O H_2S H_2Se H_2Te	HF HCl HBr HI
Formula of oxide	Li_2O Na_2O K_2O	MgO CaO SrO BaO	B_2O_3 Al_2O_3 Ga_2O_3 In_2O_3	CO_2 SiO_2 GeO_2 SnO_2 PbO_2	N_2O_3, N_2O_5 P_4O_6, P_4O_{10} As_2O_3, As_2O_5 Sb_2O_3, Sb_2O_5 Bi_2O_3 –	 SO_3 SeO_3 TeO_3 –	– Cl_2O_7 – –

The first element of each of the groups namely lithium, beryllium, boron differs in many respects from the other members of their respective group. For exampl, the behaviour of lithium is more similar to that of magnesium. This kind of similarity in properties of certain elements which are placed diagonally across in the periodic table is known as diagonal relationship. Diagonal relationships occur because of the directions in the trends of various properties as you move across or down the periodic table.

Many of the chemical properties of an element are related to the **size of the atom.**

Property	Element		
Metallic radius M/ pm	**Li** 152 **Na** 186	**Be** 111 **Mg** 160	**B** 88 **Al** 143
Ionic radius M^+ / pm	**Li** 76 **Na** 102	**Be** 31 **Mg** 72	

The anomalous behaviour is attributed to their small size, large charge/ radius ratio and high electronegativity of the elements.

Summary

Mendeleev's periodic law is based on atomic masses of the elements. The modern periodic law proposed by Mosley recognizes the properties of the elements as the periodic functions of their atomic numbers. Elements are classified into s,p(Representative), d(transition) and f (inner transition) on the basis of the modern periodic table. Elements following uranium are called transuranic elements. Ionisation enthalpy, electron gain enthalpy, and electronegativity are some of the periodic properties which regularly decreases down the group and increases along the period. Atomic and ionic radii are the periodic properties which increase down the group and decreases along the period. According to the modern periodic table, elements on the extreme left form basic oxides, those at the right forms acidic oxides and amphoteric or neutral oxides are formed by the elements in the middle.

Some Important FAQs Related to Periodic Table

1. Which law states that the atomic mass of the central element is the arithmetic mean of the atomic mass of the other 2 elements?
A. Newland’s law of octaves
B. **Doberiner’s law of triads**
C. Mendeleev periodic table
D. Modern periodic table

2. In the Modern Periodic Table, there are
A. 18 Groups and 09 periods
B. **18 Groups and 07 periods**
C. 07 Groups and 18 periods
D. 08 Groups and 06 periods

3. In Moseley’s periodic table elements are arranged according to-
A. **Increasing atomic number**
B. Increasing atomic weight
C. Increasing reactivity
D. Types of element

4. Which elements in the periodic table are chemically inactive?
A. Representative elements
B. **Inert gases**
C. Transitional elements
D. Inner transitional elements

5. Which of the following atomic number falls in the 2nd group and the 4th period of the periodic table?
A. **20**
B. 22
C. 18
D. 10

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6. In a periodic table, while moving from left to right in a period, the number of ________ remains the same.
A. electrons
B. protons
C. **shells**
D. All of these

7. In the periodic table of elements, on moving from left to right across a period, the atomic radius_______.
A. **Decreases**
B. Increases
C. Remains unchanged
D. Does not follow a definite pattern

8. The modern periodic table consists of 18 groups and 7 periods. What is the atomic number of the element placed in the 1^{st} group and the 4^{th} period?
A. 21
B. 17
C. **19**
D. 9
9. The modern periodic table consists of 18 groups and 7 periods. What is the atomic number of the element which is kept in the 4th row and 4th period?
A. 24
B. 20
C. **22**
D. 12

10. Which groups of the periodic table are known as s-block elements?
A. **1 and 2**
B. 3 - 12
C. 13 - 18
D. None of these

11. Which block elements of the periodic table are also known as inner transition elements?
A. d-block
B. **f-block**
C. s-block
D. p-block

12. Which of these elements in the periodic table are also called semi-metals?
A. Metals
B. Non Metals
C. **Metalloids**
D. Gases

13.Electronic configuration of an element is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^1$, $3d^5$. The block and period of the element in the periodic table will be
A. p block, IInd period
B. d block, IIIrd period
C. **d block, IVth period**

D. s block, IIIrd period

14. Modern periodic law had been given by ____.
A. **Moseley**
B. Mendeleev
C. Lother-Mayer
D. Lavoisier

15. Who is known as the father of the periodic table?
A. **Dmitri Mendeleev**
B. Antoine Lavoisier
C. John Newlands
D. Henry Moseley

16. How many elements are there in the 5^{th} period of the modern periodic table?
A. 2
B. 8
C. **18**
D. 36

17.The elements of a group in the periodic table
A. **have similar chemical properties.**
B. have consecutive atomic numbers
C. are isobars
D. are isotopes

18. The modern periodic table consists of 18 groups and 7 periods. What is the atomic number of the element placed in the 3rd group and the 4th period?
A. 23
B. **21**
C. 19
D. 11

19. What is the common characteristic of the elements of the same group in the periodic table?
A. **Electrons in outer most shell**
B. Total number of electrons
C. Total number of protons
D. Atomic weight

20. Which metal is the heaviest in the periodic table among the following?
A. **Os**
B. Pt
C. Ca
D. W

21. Which one of the following is the basis of the modern periodic table?
A. Atomic mass
B. **Atomic number**
C. Atomic size
D. Atomic volume

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22. Which one of the following is not a periodic property i.e. does not show any trend on moving from one side to the other in the periodic table?

A. Atomic size

B. Valency

C. **Radioactivity**

D. Electronegativity