Course Name: Chemistry
Course NO: CHE1203

Periodic Table

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Electronic configuration of elements

#	Element	Electron configuration
1	Hydrogen	$1s^1$
2	Helium	$1s^2$
3	Lithium	$1s^2 2s^1$
4	Beryllium	$1s^2 2s^2$
5	Boron	$1s^2 2s^2 2p^1$
6	Carbon	$1s^2 2s^2 2p^2$
7	Nitrogen	$1s^2 2s^2 2p^3$
8	Oxygen	$1s^2 2s^2 2p^4$
9	Fluorine	$1s^2 2s^2 2p^5$
10	Neon	$1s^2 2s^2 2p^6$

11	Sodium	$1s^2 2s^2 2p^6 3s^1$
12	Magnesium	$1s^2 2s^2 2p^6 3s^2$
13	Aluminum	$1s^2 2s^2 2p^6 3s^2 3p^1$
14	Silicon	$1s^2 2s^2 2p^6 3s^2 3p^2$
15	Phosphorous	$1s^2 2s^2 2p^6 3s^2 3p^3$
16	Sulfur	$1s^2 2s^2 2p^6 3s^2 3p^4$
17	Chlorine	$1s^2 2s^2 2p^6 3s^2 3p^5$
18	Argon	$1s^2 2s^2 2p^6 3s^2 3p^6$
19	Potassium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
20	Calcium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
21	Scandium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
22	Titanium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
23	Vanadium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
24	Chromium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
25	Manganese	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
26	Iron	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
27	Cobalt	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$
28	Nickel	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
29	Copper*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
30	Zinc	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

Using abbreviation the electronic configuration can be presented as

Z	Symbol	Electronic Configuration	Another way
1	H	$1s^1$	K1
2	He	$1s^2$	K2
3	Li	[He] $2s^1$	K2 L1
4	Be	[He] $2s^2$	K2 L2
5	B	[He] $2s^2\ 2p^1$	K2 L3
6	C	[He] $2s^2\ 2p^2$	K2 L4
7	N	[He] $2s^2\ 2p^3$	K2 L5
8	O	[He] $2s^2\ 2p^4$	K2 L6
9	F	[He] $2s^2\ 2p^5$	K2 L7
10	Ne	[He] $2s^2\ 2p^6$	K2 L8
11	Na	[Ne] $3s^1$	K2 L8 M1
12	Mg	[Ne] $3s^2$	K2 L8 M2
13	Al	[Ne] $3s^2\ 3p^1$	K2 L8 M3
14	Si	[Ne] $3s^2\ 3p^2$	K2 L8 M4
15	P	[Ne] $3s^2\ 3p^3$	K2 L8 M5
16	S	[Ne]$3s^2\ 3p^4$	K2 L8 M6
17	Cl	[Ne] $3s^2\ 3p^5$	K2 L8 M7
18	Ar	[Ne] $3s^2\ 3p^6$	K2 L8 M8
19	K	[Ar] $4s^1$	K2 L8 M8 N1

Periodic Table

Mendeleeff's Periodic Law

The properties of elements are a periodic function of their atomic weights, i.e. if the elements are arranged in the increasing order of their atomic weights, the properties of the elements are repeated after definite regular intervals or periods.

Defects:

(i) Position of hydrogen

ii) Position of lanthanides and actinides

A group of 15 elements (57 to 71) does not find its proper place in the table, is in g III and period 6.

Similarly, another group of 15 elements (89 to 103) is in group III and period 7.

(iii) **Similar elements are separated while dissimilar elements are placed in the same group.**
Separated similar elements: Cu, Ag; Au, Tl; Ba, Pb
Dissimilar elements grouped together: Cu, Ag and Au are grouped along with the alkali metals.

(iv) **Existence of four anomalous pair of elements**
The elements of higher atomic weight precede those of lower atomic weight at four places as shown below:

(a) Ar(Z=18, at.wt.=40) precedes K(Z=19,at.wt.=39)
(b) Co(Z=27, at.wt.=59) precedes Ni(Z=28,at.wt.=58)
(c) Te (Z=52, at.wt.=127.6) precedes I(Z=53,at.wt.=126.9)
(d) Th (Z=90, at.wt.=232) precedes Pa(Z=91,at.wt.=231)

(v) Position of isotopes

(vi) Group does not represent valency

For example, Os placed in group VIII does not show a valency of 8.

Modern Periodic law

The modern periodic law states that the physical and chemical properties of the elements are the periodic functions of their atomic numbers, i.e. if the elements are arranged in the increasing order of their atomic numbers, the properties of the elements are repeated after definite regular intervals or periods.

Mendeleeff's Irregularities Disappear

(i) **Position of hydrogen**

The dual role of hydrogen is explained by the fact that it has one electron in its outer orbit which is being placed up and replaceable as H^- and H^+ is in group VII and I.

(ii) **Anomalous pairs of elements**

This anomaly disappears altogether and the pairs Ar-K, Co-Ni, Te-I and Th-Pa and are found in the table in the order of increasing atomic number.

Pairs of elements:	Ar	K	Co	Ni	Te	I	Th	Pa
Atomic No.	18	19	27	28	52	53	90	91
Atomic wt.	40	39	59.9	58.6	127.6	126.7	232	231

(iii) Position of rare earth metals

(iv) Position of isotopes

(v) Justification of dissimilar elements being placed together.

The Modern Periodic Table

- First short period contains only two elements (H, He) $1s^2$
- The second short period contains eight elements beginning with Li and ending with Ne. Neon has complete electronic configuration of $1s^2 2s^2 2p^6$ and contains 8 electrons in valence shell.
- The third short period again consists of 8 elements (Na-Ar). Ar has electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^6$.
- The fourth period (1st long period) contains 18 elements (K-Zn). It is ending with Krypton containing 8 electron in the outermost shell.

- The fifth period begins with rubidium, Rb(37). The last element is xenon, Xe(54) and contains 8 electrons in the outermost shell.
- The sixth period consists of 32 elements. It starts from cesium, Cs(55) and next is Ba with $6s^2$, the next 15 elements La(57) to Lu(71). The 15 elements La(57) to Lu(71) have almost identical chemical properties and all are placed in the same position. They are called rare earth metals and constitutes lanthanides series. This period ends with radon.

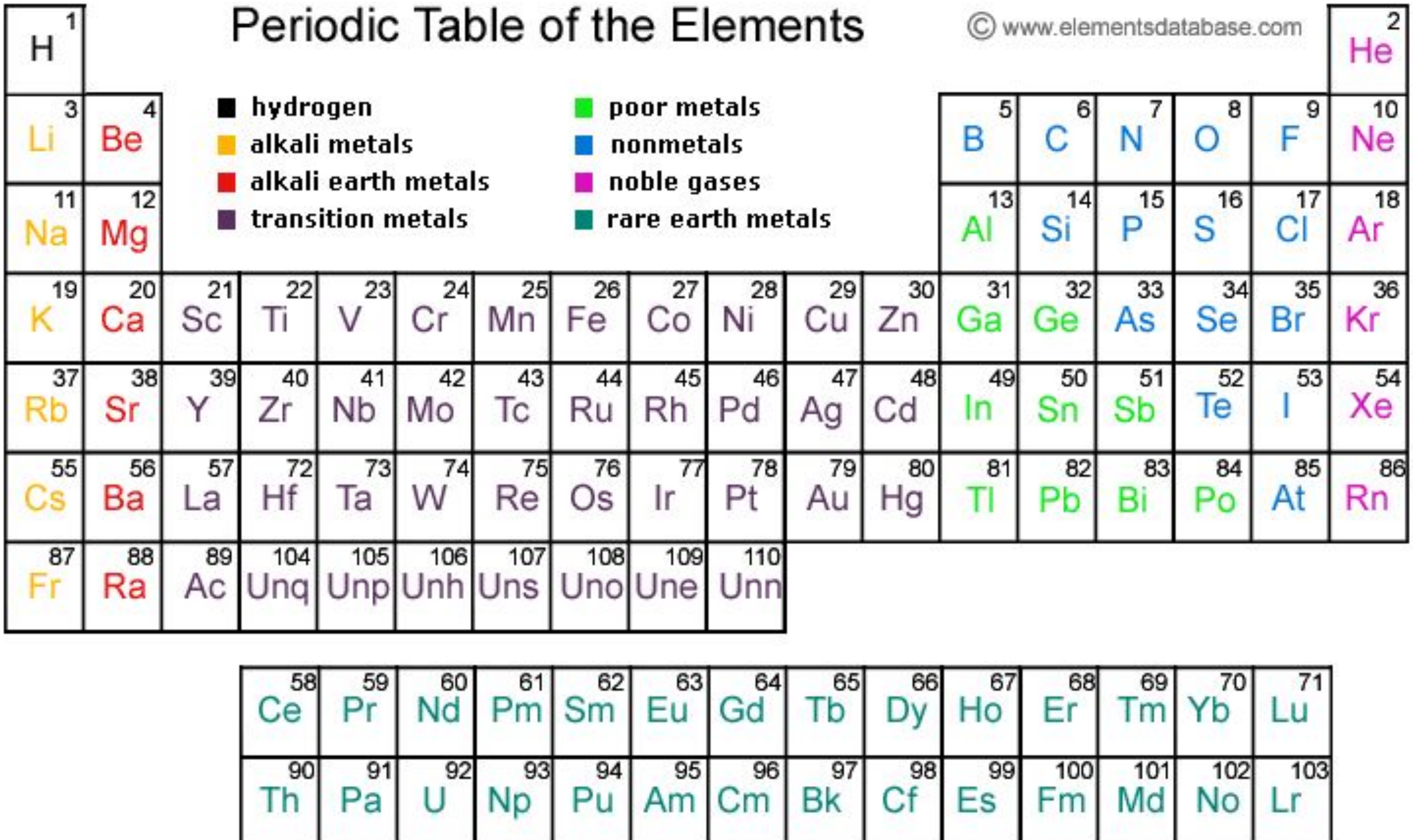
The vertical arrangement of elements in the periodic table are called group, IA, IIA, IIIB, IVB, VB, VIB, VIIB, VIII, IB, IIB, IIIA, IVB, VA, VIA,VIIA, 0.

The elements in the sub group A shows same similarities to the corresponding elements in the B sub group of the same number.

The division of the sub-groups A and B is based on the fact that the penultimate energy level (last but one quantum number level) of electron in these groups contain the arrangement s^2p^2 and $s^2p^2d^{10}$ respectively.

Group VIII have no sub-groups, instead these consist of three elements in one single group of the periodic table. The group contains Fe, Co, Ni in the fourth period, ruthenium, rhodium and palladium in the fifth period, and osmium, iridium and platinum in the sixth period. They are called bridge elements.

Periodic Table



Electronic Structure and periodic law

According to electronic configurations, the elements may be divided into four types.

(a) The inert gases (Elements of ‘0’ group)-ns^2np^6

(b) The representative elements (s and p block elements)-ns^1, ns^2, ns^2np^{1-5}.

(c) The transition elements (d block elements)-$(n-1)d^{1-10}ns^2$

(d) The inner transition elements (f block elements)-three incomplete level

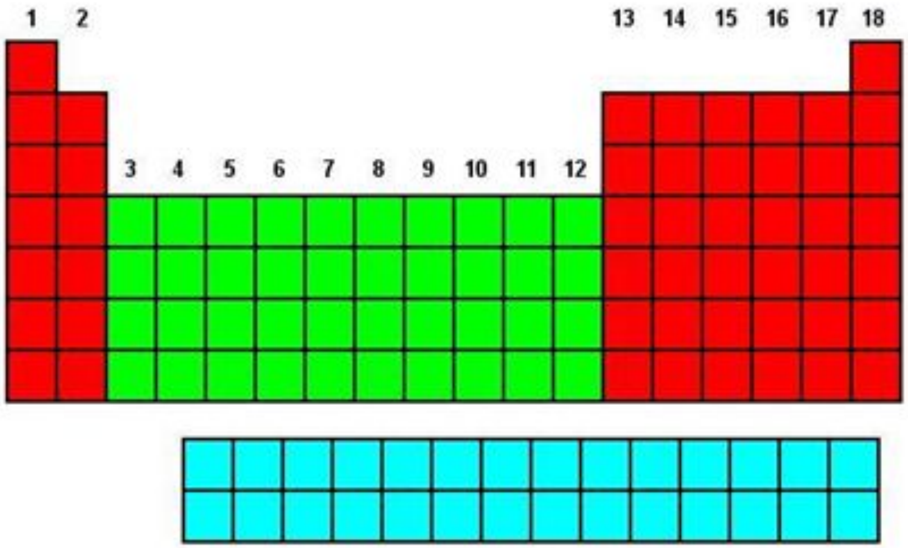
Inert gases

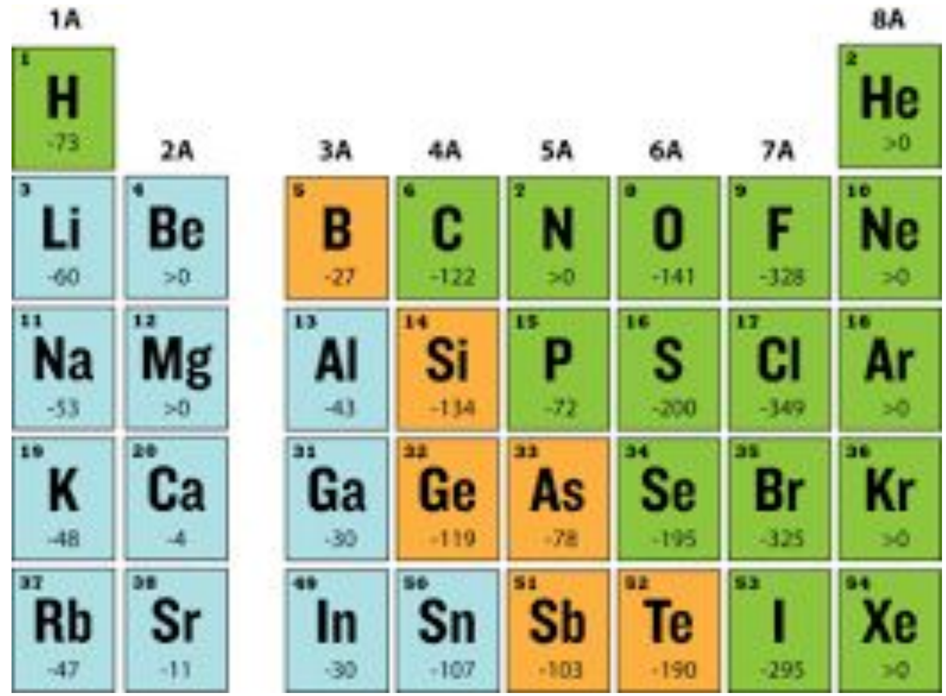
Complete outermost electronic configuration: ns^2np^6.

Element	Atomic Number	Electronic Configuration
Helium	2	$1s^2$
Neon	10	$1s^22s^22p^6$
Argon	18	$1s^22s^22p^63s^23p^6$
Krypton	36	$1s^22s^22p^63s^23p^63d^{10}4s^24p^6$
Xenon	54	$1s^22s^22p^63s^23p^63d^{10}4s^2\,4p^64d^{10}5s^25p^6$
Radon	86	$1s^22s^22p^63s^23p^63d^{10}4s^24p^6\,4d^{10}4f^{14}5s^25p^65d^{10}6s^26p^6$

The representative elements

These elements belong to A sub-group of the periodic table. They are metals and non-metals.





Periodic Table

Representative elements

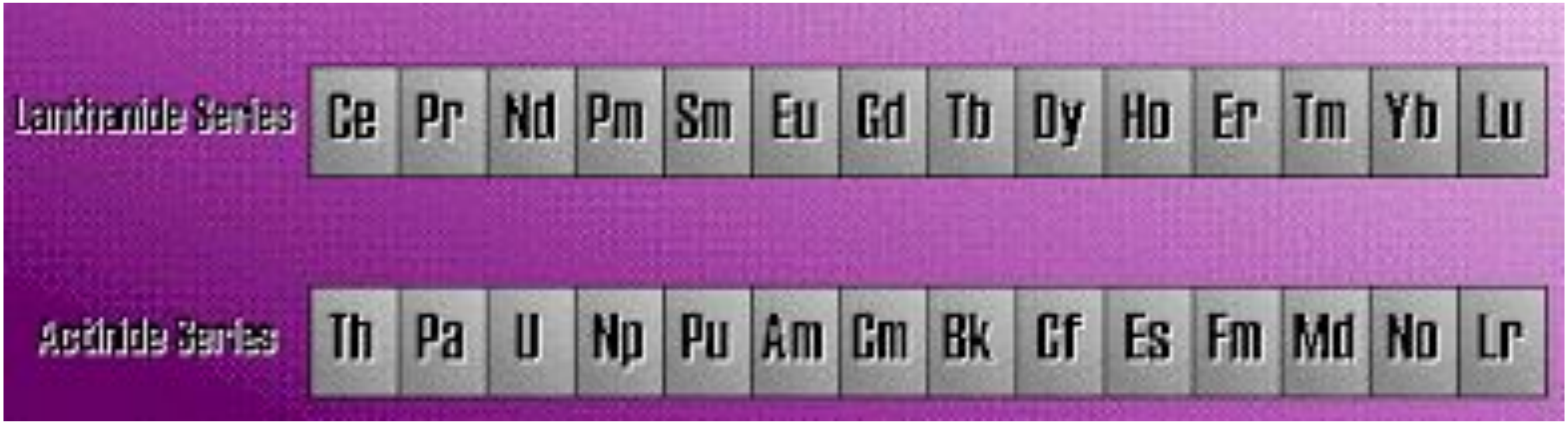
The Transition Metals

They are metals of sub-group B and containing two incomplete energy levels.

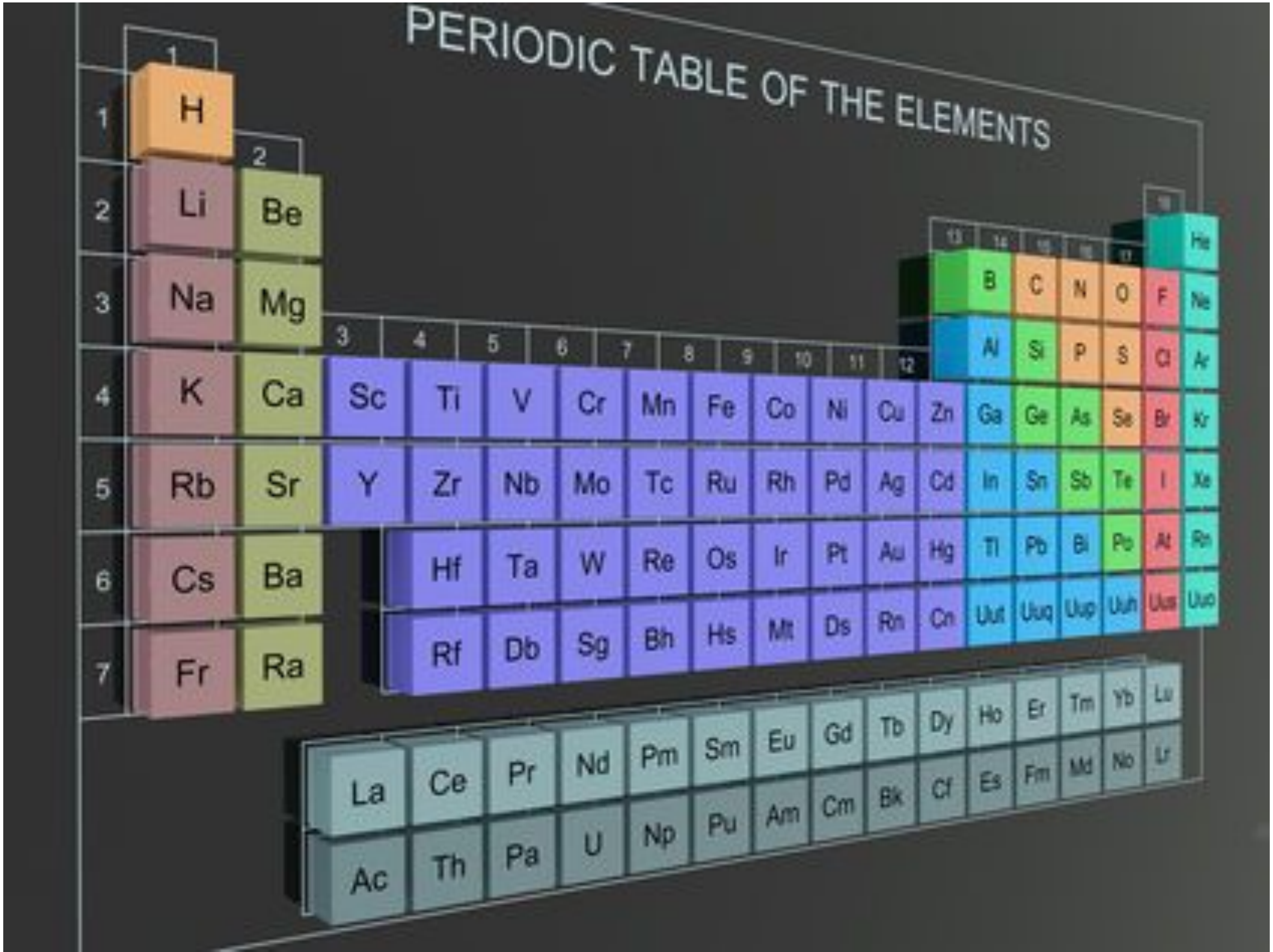
21 Sc 44.9559 Scandium	22 Ti 47.867 Titanium	23 V 50.9415 Vanadium	24 Cr 51.9961 Chromium	25 Mn 54.938 Manganese	26 Fe 55.845 Iron	27 Co 58.9332 Cobalt	28 Ni 58.6934 Nickel	29 Cu 63.546 Copper	30 Zn 65.4089 Zinc
39 Y 88.9058 Yttrium	40 Zr 91.224 Zirconium	41 Nb 92.9064 Niobium	42 Mo 85.94 Molybdenum	43 Tc 98 Technetium	44 Ru 101.07 Ruthenium	45 Rh 102.9055 Rhodium	46 Pd 106.42 Palladium	47 Ag 107.8682 Silver	48 Cd 112.411 Cadmium
71 Lu 174.967 Lutetium	72 Hf 178.49 Hafnium	73 Ta 180.9497 Tantalum	74 W 183.84 Tungsten	75 Re 186.207 Rhenium	76 Os 190.23 Osmium	77 Ir 192.217 Iridium	78 Pt 195.084 Platinum	79 Au 196.9666 Gold	80 Hg 200.59 Mercury

Inner Transition Elements

These elements have three incomplete shells. The series of 14 elements in which 4f levels is being built up follows lanthanum and are called lanthanides. The series of 14 elements in which 5f levels is being filled follows actinium known actinides.



Complete Periodic Table



Variation of properties within periods and groups

Differences in the properties of the elements from elements in a period and resemblance within groups may be attributed primarily to the characteristics of atomic structures.

(i) The nuclear charge and the number of electrons surrounding the nucleus.

(ii) The total number of electrons-particularly the number of valance electrons.

(iii) The size of the atoms, i.e. the volume occupied by the electron in various energy levels.

Variations of Metallic Character of the Elements

In the periodic table, the metallic character of the elements decreases from left to right progressing in the series but increases in moving vertically from top to bottom in the groups. Metallic character means-electrical and thermal conductivity, metallic luster, reducing properties.

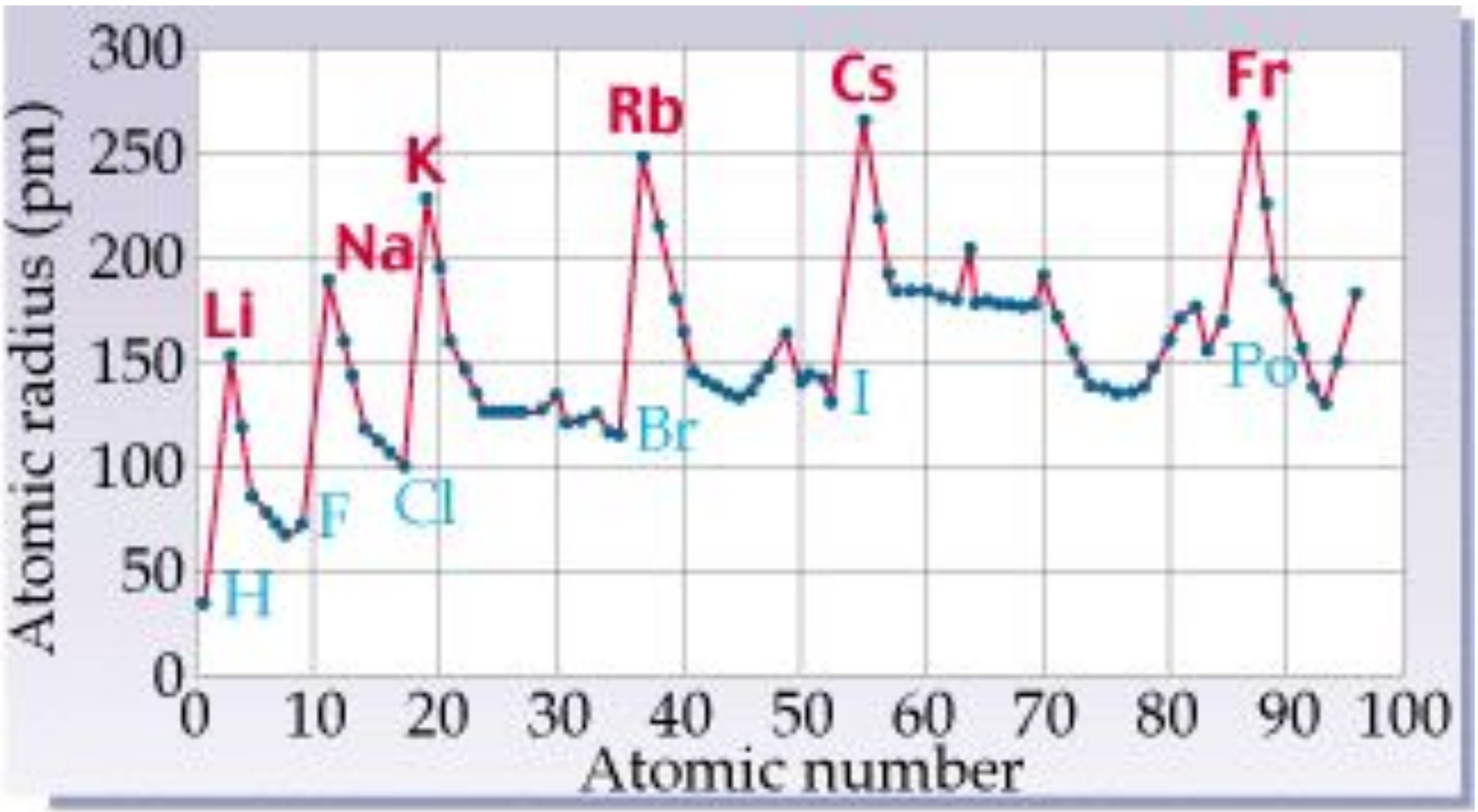
Variation of Atomic Sizes (Radii)

The atomic size of each succeeding elements in a period decreases but atomic radii of inert gases at the end of the periods, however, are larger than those of the elements of the preceding atomic number.

The inert gases have completed outer energy levels, i.e. the p sub-shell has been filled up with six electrons. As the nuclear charge increases in a period the electron may occupy the same or different energy levels. When the succeeding electrons go into the same energy levels they are subjected to greater attraction by the increased nuclear charge and hence the elements in a series show gradual decreases in the atomic size.

Vertically in the groups the succeeding elements have increasing atomic radii. Thus each alkali metal atom has a much greater atomic radius than that of the inert gas just before it.

This is due to the fact that the additional electrons occupies a new sub-shell with quantum number higher than those of the already filled energy levels.

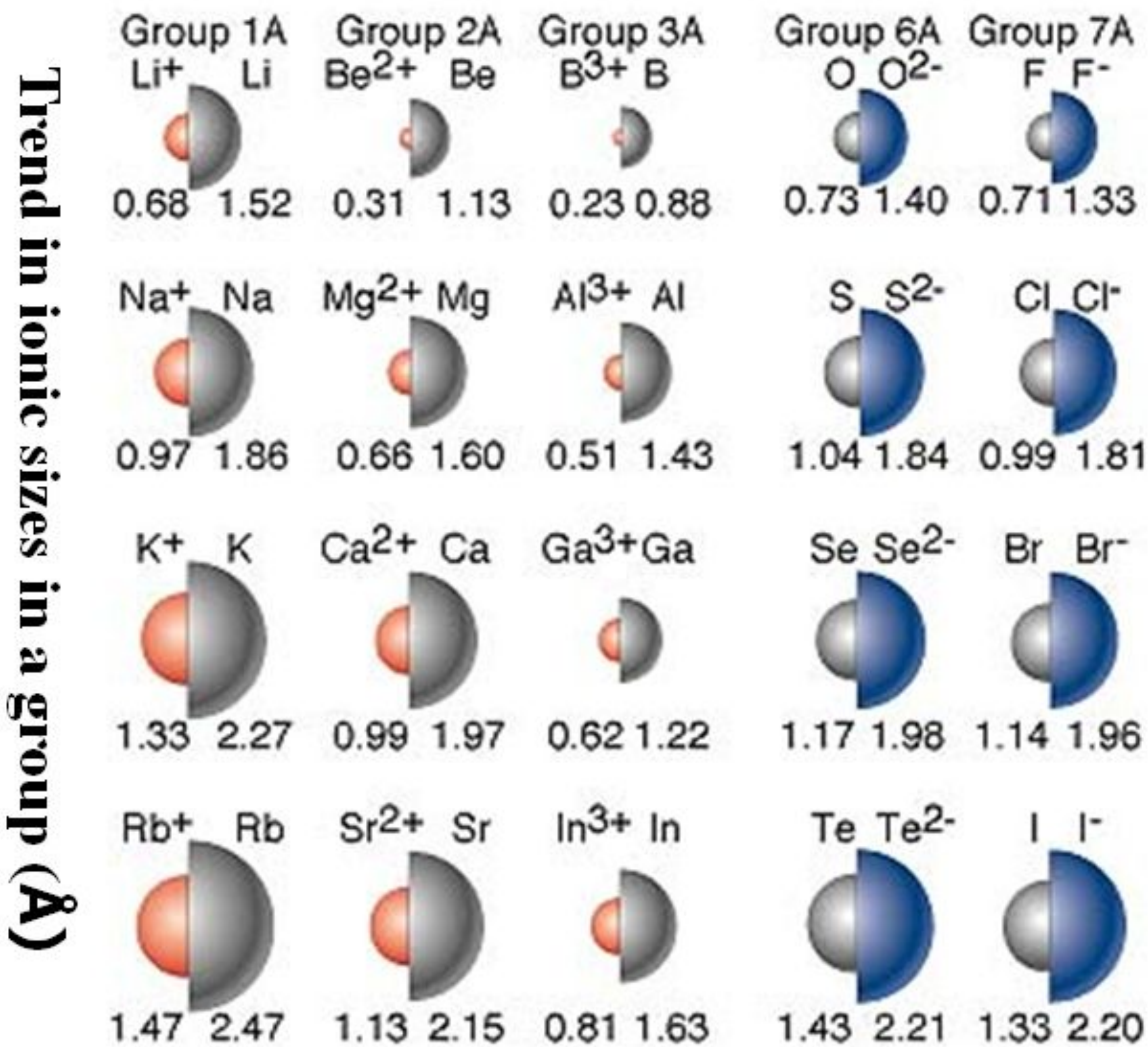


Variation of Ionic Radii

It is obvious that the size of a positive ion will be less than that of atom from which it is formed. There is considerable decrease in size due to the loss of the outermost electron particularly in the case of alkali metals. Ions with inert gas electronic configurations, the contraction in size in a given period is well-marked. Thus in the series (3rd), Na^+, Mg^{2+}, Al^{3+} and Si^{4+}, which are isoelectronic with Ne configuration, the decrease in their ionic size appears to be considerable as compared to the atomic size of the parent atom. The greater the nuclear charge, the smallest is the ionic radius in a series of isoelectronic ions.

In a given group, positive ions of succeeding elements have larger ionic radii.

Trend in ionic sizes in a series (Å)



Variation of Ionization Potentials

The ionization potential (IP) or energy of an atom is the minimum amount of energy needed to remove the highest-energy electron from the neutral atom in the gaseous state. In the case of Li atom, the 1st ionization energy is 520 KJ/mol.

$$Li\ (1s^2 2s^1) \rightarrow Li^+\ (1s^2) + e^-$$

IP display a periodic variation when plotted against atomic numbers. Within any period, values tend to increase with atomic number. Lowest value in a period are found for Gr. IA. The IP in a period occur for the inert gas elements because the inert gas atom loses electron with difficulty.

In general, the greater the nuclear charge of atoms having the same number of electron orbit, the greater the IP. Thus the elements in the same period have gradually increasing IP. Slight irregularity within a period is due to building up of new sub-levels for electrons. Thus IP increases in a series and shows decreasing tendency with a group in the periodic classification.

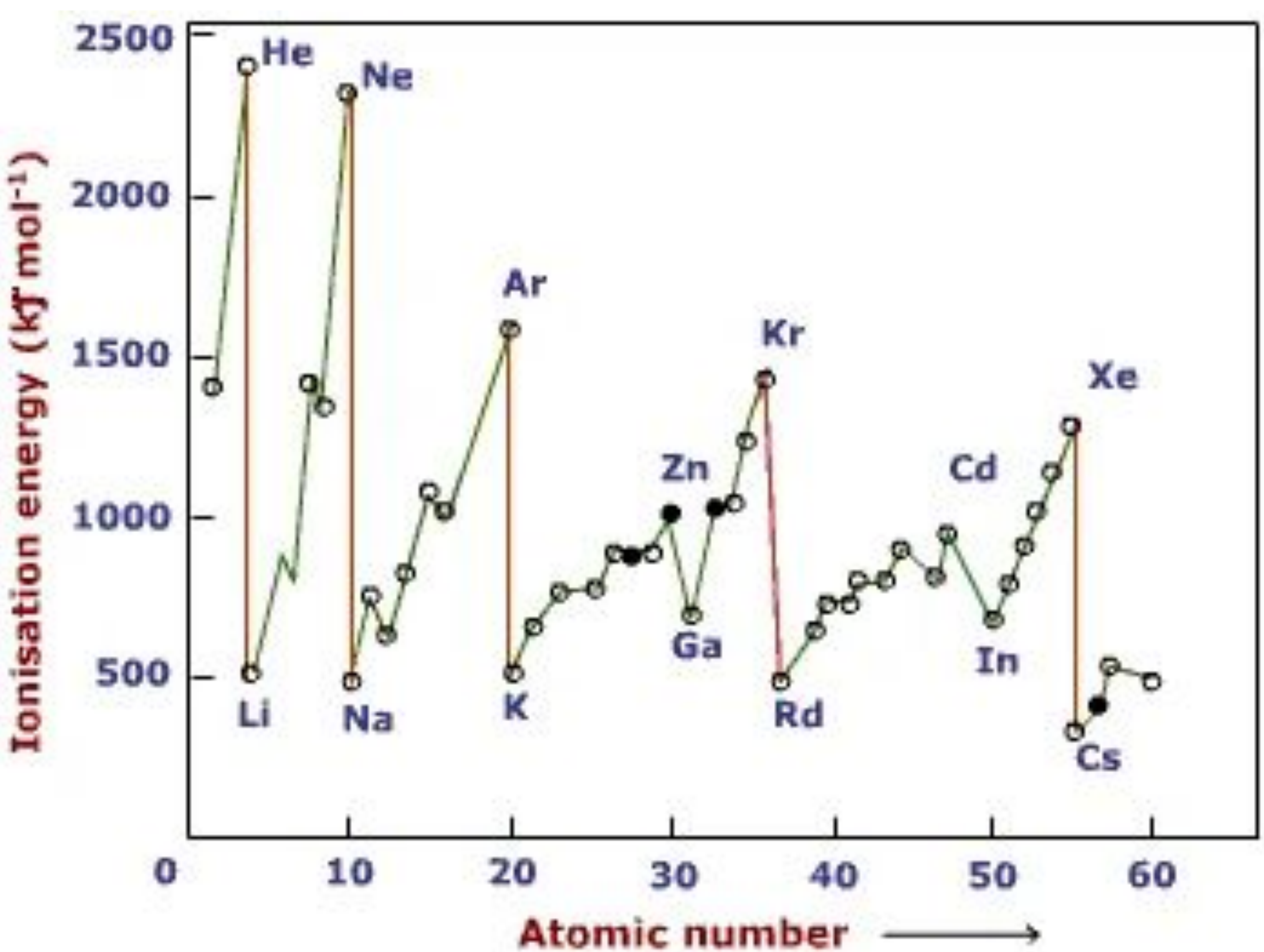


Fig. Ionization energy vs. Atomic number

Variation of Electron Affinities

The electron affinity is the energy change for the process of adding an electron to a neutral atom in the gaseous state to form a negative ion. For example, a chlorine atom can pick up an electron to give a chloride ion Cl^-, and 394 KJ/mol of energy is released.

$$Cl\ ([Ne]3s^23p^5) + e- \longrightarrow Cl-\ ([Ne]3s^23p^6)$$

Electron affinities have a periodic variation, just as atomic radii and ionization energies do.

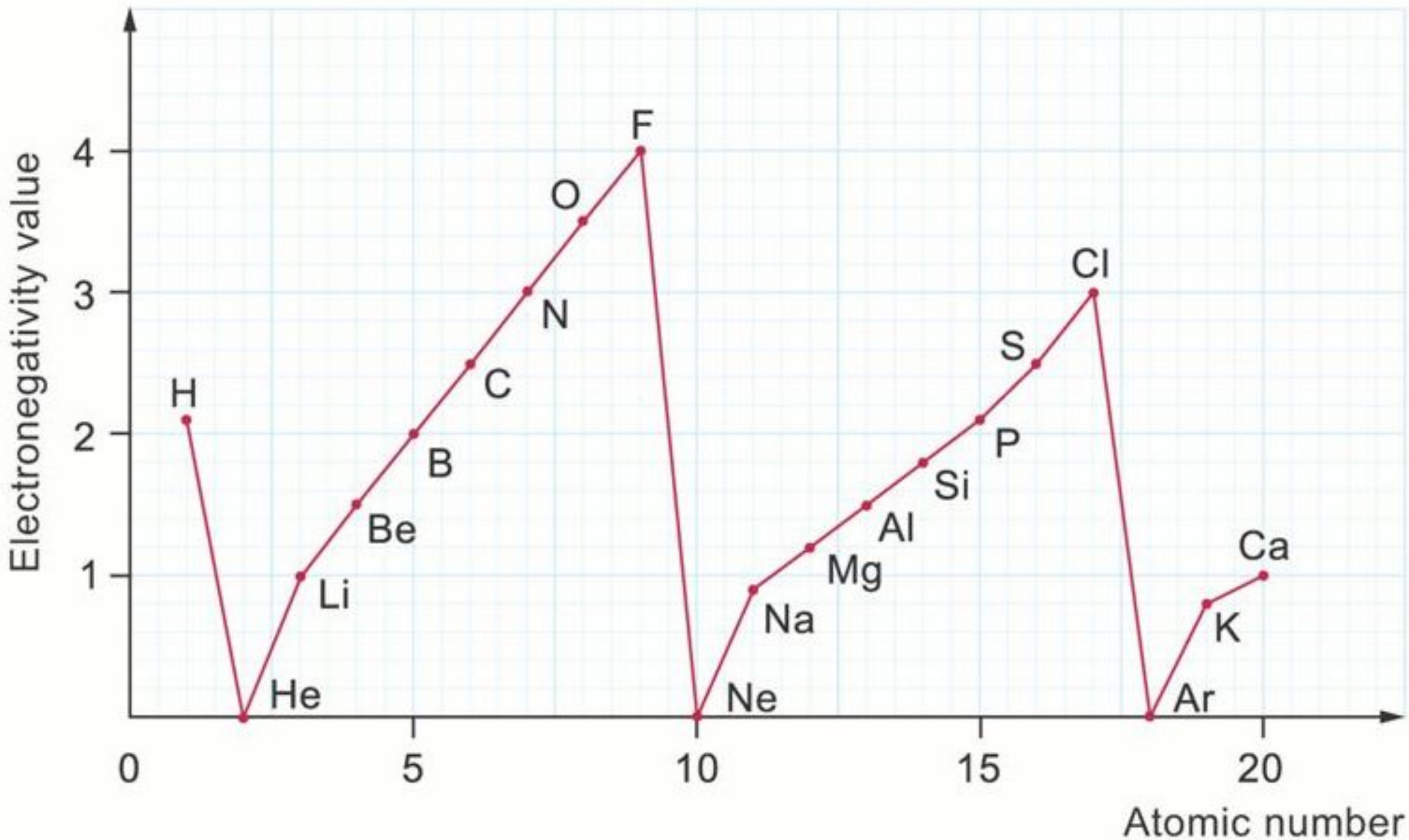
H -73	Electron Affinity s & p-block Elements kJ mol^{-1}						He -48
Li -60	Be +66	B -27	C -122	N +31	O -141	F -328	Ne +116
Na -53	Mg +67	Al -43	Si -134	P -72	S -200	Cl -349	Ar +196
K -48	Ca	Ga -29	Ge -116	As -77	Se -195	Br -324	Kr +96
Rb -47	Sr	In -29	Sn -120	Sb -101	Te -190	I -295	Xe +77
Cs -45	Ba	Tl -29	Pb -35	Bi -45	Po -183	At -270	Rn +68

Fig. Variation of electron affinity within period and along group of the periodic table

Variation of Electronegativity

The power of the attraction that an atom shows for electron in a covalent bond also shows periodic variation. Electronegativity has been defined as the power of attraction by an atom for electron in a molecule.

The most electronegative elements are found towards the end of the periods. Metals having low electronegativities are found at the beginning of the periods. Thus alkali metal show gradually decreasing electronegativity values within the group. The halogens are most electronegative elements and the values decrease from fluorine to iodine.



Variation in electronegativity values of the first 20 elements

Diagonal Relationship

Some elements of certain groups of second period resemble much in properties with the elements of third period of next group i.e. elements of second and third period are diagonally related in properties. This phenomenon is known as diagonal relationship.

Period	Group			
	IA	IIA	IIIA	IVA
2	Li	Be	B	C
3	Na	Mg	Al	Si

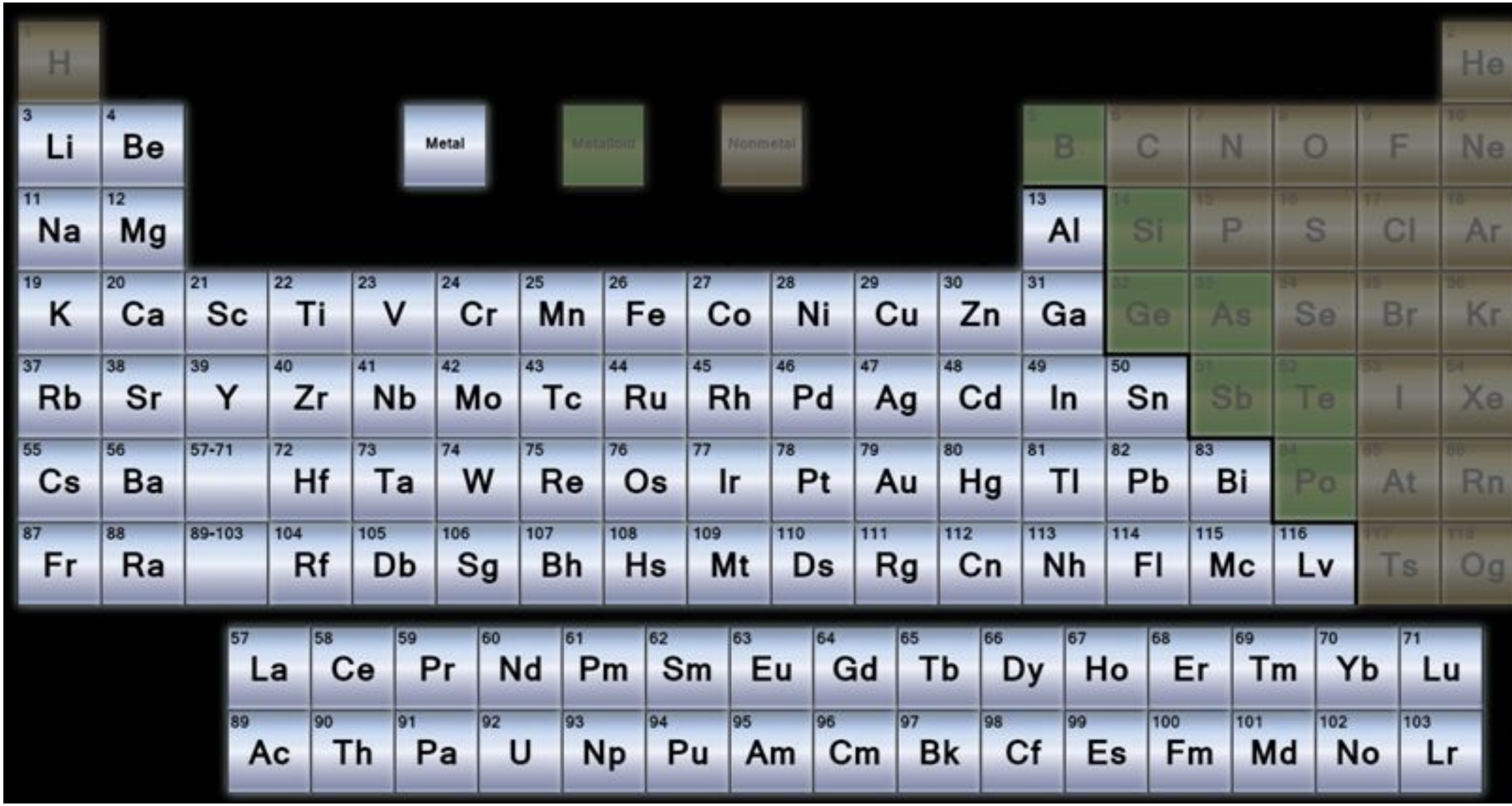
Lithium, a member of the alkali metal of Gr. IA, in some respects resembles magnesium of Gr. IIA. Thus unlike other alkali metal carbonates and phosphates Li_2CO_3 and Li_3PO_4 are insoluble in water, as are the corresponding $MgCO_3$ and $Mg_3(PO_4)_2$. Li is the only alkali metal to form Li_3N, like magnesium nitride. Thus Li of Gr. IA resembles Mg of Gr.IIA in many respects contrary to its group properties. Similar relationship exists between the three elements, Be of Gr. IIA, Al of Gr. IIIA, B of Gr. IIIA shows likeness with Si of Gr. IVA. Thus the light elements of one group shows similarity in properties with second elements of the following groups.

This similarity is generally referred to as diagonal relationship in the periodic table.

Metals

A metal is a material that, when freshly prepared, polished, or fractured, shows a lustrous appearance, and conducts electricity and heat relatively well. Metals are typically malleable (they can be hammered into thin sheets) or ductile (can be drawn into wires). A metal may be a chemical element such as iron; copper, cobalt. Metal can easily release electron.

$$M \rightarrow M^+ + e^-$$



1 H
2 He
Metal
Metalloid
Nonmetal
3 Li
4 Be
5 B
6 C
7 N
8 O
9 F
10 Ne
11 Na
12 Mg
13 Al
14 Si
15 P
16 S
17 Cl
18 Ar
19 K
20 Ca
21 Sc
22 Ti
23 V
24 Cr
25 Mn
26 Fe
27 Co
28 Ni
29 Cu
30 Zn
31 Ga
32 Ge
33 As
34 Se
35 Br
36 Kr
37 Rb
38 Sr
39 Y
40 Zr
41 Nb
42 Mo
43 Tc
44 Ru
45 Rh
46 Pd
47 Ag
48 Cd
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
55 Cs
56 Ba
57-71
72 Hf
73 Ta
74 W
75 Re
76 Os
77 Ir
78 Pt
79 Au
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
87 Fr
88 Ra
89-103
104 Rf
105 Db
106 Sg
107 Bh
108 Hs
109 Mt
110 Ds
111 Rg
112 Cn
113 Nh
114 Fl
115 Mc
116 Lv
117 Ts
118 Og
57 La
58 Ce
59 Pr
60 Nd
61 Pm
62 Sm
63 Eu
64 Gd
65 Tb
66 Dy
67 Ho
68 Er
69 Tm
70 Yb
71 Lu
89 Ac
90 Th
91 Pa
92 U
93 Np
94 Pu
95 Am
96 Cm
97 Bk
98 Cf
99 Es
100 Fm
101 Md
102 No
103 Lr

Non-Metal

Non-metals are those which lack all the metallic attributes. They are good insulators of heat and electricity. They are mostly gases and sometimes liquid. Some they are even solid at room temperatures like Carbon, sulphur and phosphorus. Nonmetal can accept electron easily. $X + e^{-} \rightarrow X^{-}$

Properties of Non-metals

Characteristics properties of non-metals are high ionization energies and high electronegativity. Because of these properties non-metals usually gain electrons when reacting with other compounds, forming covalent bonds.

The following are the general properties of non-metals.

(i) The atoms of non-metals tend to be smaller than those of metals. Several of the other properties of non-metals result from their atomic sizes.

(ii) Non-metals exhibit very low electrical conductivities. The low or non-existent electrical conductivity is the most important property that distinguishes non-metals from metals.

(iv) Non-metals have high electronegativities. This means that the atoms of non-metals have a strong tendency to hold on to the electrons that already have.

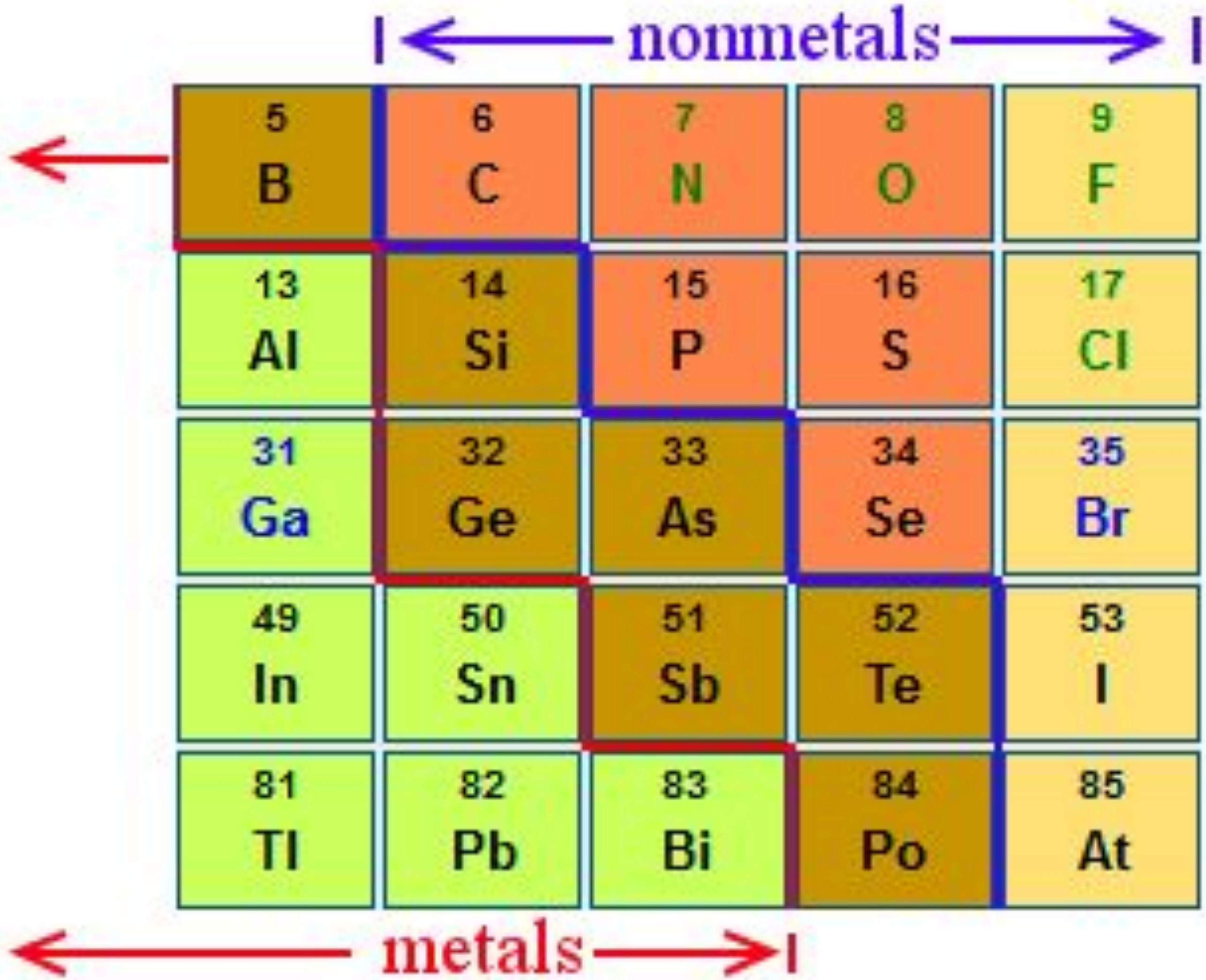
In contrast, metals rather easily give up one or more electrons to non-metals, metal therefore easily form positively charged ions, and metals readily conduct electricity.

(v) Under normal conditions of temperature and pressure, some non-metals are found as gases, some found as solids and one is found as liquid. In contrast, except mercury, all metals are solids at room temperature. The fact that so many non-metals exist as liquids or gases means that non-metals generally have relatively low melting and boiling points under normal atmospheric conditions.

(vi) In their solid-state, non-metals tend to be brittle. Therefore, they lack the malleability and ductility exhibited by metals.

Metalloid

A metalloid is a chemical element that exhibits some properties of metals and some of nonmetals. In the periodic table metalloids form a jagged zone dividing elements that have clear metallic properties from elements that have clear nonmetallic properties. Boron, silicon, germanium, arsenic, antimony, tellurium, and polonium are metalloids. In some cases, authors may also class selenium, astatine, aluminum, and carbon as metalloids, but this is less common.


nonmetals
5 B
6 C
7 N
8 O
9 F
13 Al
14 Si
15 P
16 S
17 Cl
31 Ga
32 Ge
33 As
34 Se
35 Br
49 In
50 Sn
51 Sb
52 Te
53 I
81 Tl
82 Pb
83 Bi
84 Po
85 At
metals

Noble Gases

Noble gas, any of the seven chemical elements that make up Group 18 (Zero) of the periodic table. The elements are helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), radon (Rn) and oganesson (Og). The noble gases are colourless, odourless, tasteless, nonflammable gases. They traditionally have been labeled Group 0 in the periodic table because for decades after their discovery it was believed that they could not bond to other atoms; that is, that their atoms could not combine with those of other elements to form chemical compounds.

Their electronic structures and the finding that some of them do indeed form compounds has led to the more appropriate designation, Group 18.

	symbol	electron configuration
helium	He	$1s^2$
neon	Ne	$[He]2s^22p^6$
argon	Ar	$[Ne]3s^23p^6$
krypton	Kr	$[Ar]3d^{10}4s^24p^6$
xenon	Xe	$[Kr]4d^{10}5s^25p^6$
radon	Rn	$[Xe]4f^{14}5d^{10}6s^26p^6$

Properties of Noble gases

Physical Properties of Noble Gases

Properties	He	Ne	Ar	Kr	Xe	Rn
Ionization Energy (kjmol-1)	2372	2081	1521	1351	1170	1037
Atomic Radius(pm)	40	70	94	109	130	140
Melting Point(°C)	-272	-249	-189	-157	-112	-71
Boiling Point(°C)	-269	-246	-186	-153	-108	-61
Water Solubility (ml/lit)at 20 °C	13.8	14.7	37.9	73.00	110.9	....
Heat of vaporization (kj/mol-1)	0.08	1.77	6.5	9.7	13.7	18.0