# Periodic Table

# Mendeleev's Periodic Classification

The Periodic Law stated by Mendeleev in 1869 was the first successful attempt in the classification of elements and their comparative study. According to this law:

“The physical and chemical properties of elements are a periodic function of their atomic weights.”

# Periodic Law

With the advancement of our knowledge about atomic structure, a modification in the definition of periodic law was proposed on the basis of atomic number. The modern statement of the periodic law is:

"The physical and chemical properties of the elements are periodic function of their atomic numbers."

# The Modern Periodic Table

# Study of Periods and Groups

**Periods:** These are horizontal rows. Long form of the periodic table consists of 7 periods.

***(I) First Period (n=1).*** n=1 indicates that there is only one main energy shell for the elements of this period. This period has two elements namely $H_1$ and $He_2$.

***(II) Second Period (n=2).*** There are two shells in the elements for this period. It has 8 elements $Li_3$ to $Ne_{10}$.

***(iii) Third Period (n=3).*** This period also has 8 elements, $Na_{11}$ to $Ar_{18}$.

Second and third periods are called short periods.

***(iv) Fourth Period (n=4).*** This period has 18 elements, $K_{19}$ to $Kr_{36}$.

***(v) Fifth period (n=5).*** This period also has 18 elements, $Rb_{37}$ to $Xe_{54}$.

***(vi) Sixth period (n=6).*** This period has 32 elements, $Cs_{55}$ to $Rn_{86}$.

Fourth, fifth and sixth periods are called long periods.

***(vii) Seventh Period (n=7).*** This is an incomplete period which at present has 26 elements, $Fr_{87}$ to $Cn_{112}$.

**Groups:** The vertical columns are called groups. These are 16 in all as shown below although the total number of vertical column is 18.

***(I) IA, IIA,IIIA,IVA, VA, VIA and VIIA groups (7 groups).*** The elements of these groups are called normal elements or main group elements or representative elements.

***(II) IB, IIB, IIIB, IVB, VB, VIB, VIIB, and VIII groups (8 groups).*** Group VIII has three columns.

IB, IIB, IIIB, IVB, VB, VIB, VIIB and VIII groups have purely transition elements while in group IIIB some elements are transition elements ( e.g. $Sc_{21}$, $Y_{39}$, $La_{57}$ and $Ac_{89}$ 4 elements) while the remaining (e.g. $Ce_{58}$ to $Lu_{71}$ and $Th_{90}$ to $Lw_{103}$ 28 elements) are inner transition elements.

***(iii) Zero group.*** This group has inert gas elements.

Each long period consists of two series, *first series* and *second series* and because of these two series, we get **Sub-Group A** and **Sub-Group B**.

The elements of **Sub-Group A** resemble one another but differ from the elements of **Sub-Group B** of the same groups, all elements of the group showing the same valence.

Members of one sub-group have similar physical and chemical properties which vary gradually with increase of atomic number.

After *lanthanum*(57) are placed fourteen rare earth elements (58-71) known as *lanthanides*, very similar in properties amongst themselves.

Their compounds are so closely related to one another that their separation involved difficulties.

Similarly after Actinium (89), there is another group of fourteen elements very similar in their properties known as actinides (At. No: 90-103).

# Long Form Periodic Table

The periodic table is constructed so that elements with the same type of valence electron configuration are arranged in columns.

The *last* electron added to an atom in the building-up process, is called the **distinguishing electron** because it distinguishes an atom from the one immediately preceding it, in the periodic table.

The **representative elements** are those in which the distinguishing electron enter an *s* or *p* subshell.

The commonly used long form of the periodic table is designed to emphasize electron configurations.

# Electron Configurations and the Periodic Table

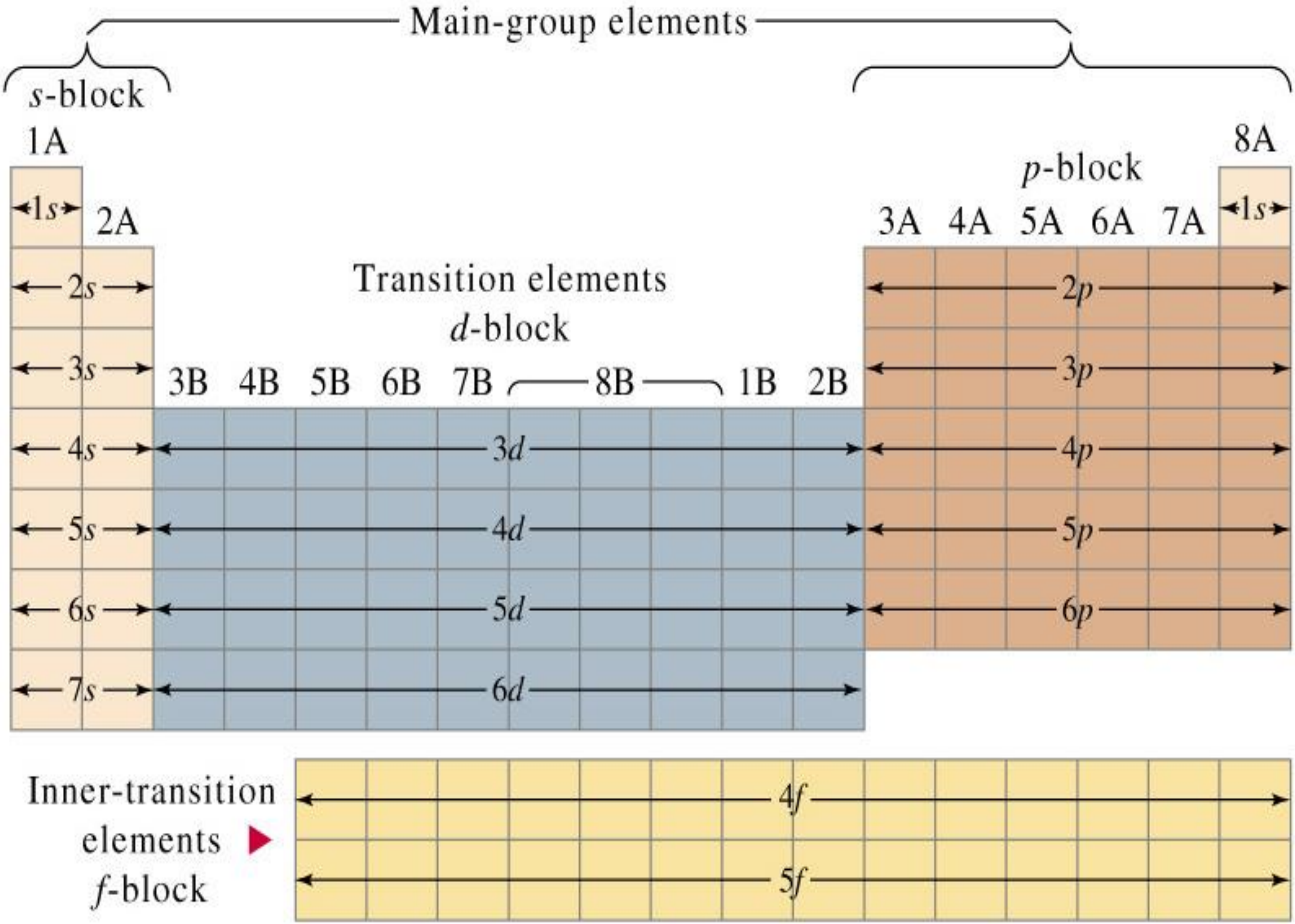
According to the electronic configurations, the elements may be divided into four types. These are:

(1) The Inert Gases. (Elements of 0 group)

(2) The Representative Elements. (s and p block elements)

(3) The Transition Elements. (d block elements)

(4) The Inner Transition Elements. (f block elements)



Main-group elements
s-block
1A
2A
p-block
8A
3A
4A
5A
6A
7A
1s
1s
2s
3s
4s
5s
6s
7s
Transition elements
d-block
3B
4B
5B
6B
7B
8B
1B
2B
3d
4d
5d
6d
2p
3p
4p
5p
6p
Inner-transition
elements
f-block
4f
5f

**(1) The Inert Gases:**

The zero group elements have been placed at the end of each period in the Periodic Table. It appears that these elements having $s^2p^6$ electronic arrangements in the outermost level are very stable. Helium has $2s^2$ stable arrangement and all other inert gases have $s^2p^6$ outer configurations.

**(2) The representative Elements (s and p block elements):**

These elements generally belong to **A** sub-group of the Periodic Table. These elements have the outermost energy level incomplete. The chemical behaviour of these elements depends upon the valence electrons.

***i. s-block Elements.*** In these elements, the differentiating electron enters the **ns** orbital which is being progressively filled. The elements of groups **IA** and **IIA** belong to this block.

***In these elements the valence s orbitals are being filled*** **.** The valence shell configuration of these elements varies from ***ns*** to $\boldsymbol{ns^2}$.

The members of *s-block* elements lie on the extreme left of the periodic table and called alkali metals and the alkaline earth metals**.**

***ii. P-block Elements.*** The elements in which p-orbitals are being progressively filled are called p-block elements.

The elements of the groups **IIIA, IVA, VA, VIA, VIIA** and **zero** (with the possible exception of helium whose configuration is $1s^2$) are the members of this block, since in the atoms of these elements, the differentiating electron enters the p-orbitals, namely, *2p, 3p, 4p, 5p* and *6p* respectively.

The *s-orbitals* in these atoms are already completely filled. Thus the valence configuration of the atoms of these elements vary from $ns^2p^1$ (group **IIIA**) to $ns^2p^6$ (**zero** group). These elements lie at the extreme right of the periodic table and consists of some metals, all non-metals, metalloids and noble gases.

## (3) Transition Elements (d-block Elements):

The elements in which the differentiating electron enters the *(n-1) d-orbital* of the *(n-1)*th main shell are called d-block elements. These elements is a block of ten columns and are placed in the middle of the periodic table, between *s*- and *p-block* elements. These elements are also called transition elements

These elements are generally heavy metals of sub-group **B** and contain two incomplete energy levels because of the building up of the inner *d* electrons. The chemical properties of these elements depends upon the electrons from the two outermost levels (*s* and *d* electrons).

The elements in group VIII have no sub-groups, instead these consist of three elements in one single group of the Periodic Table and occur in the middle of each long period. The group VIII elements may be called bridge elements, because they form bridge between the hard, high melting metals of group VIB and VIIB and the softer metals of group IB and IIB.

These elements generally form coloured compounds.

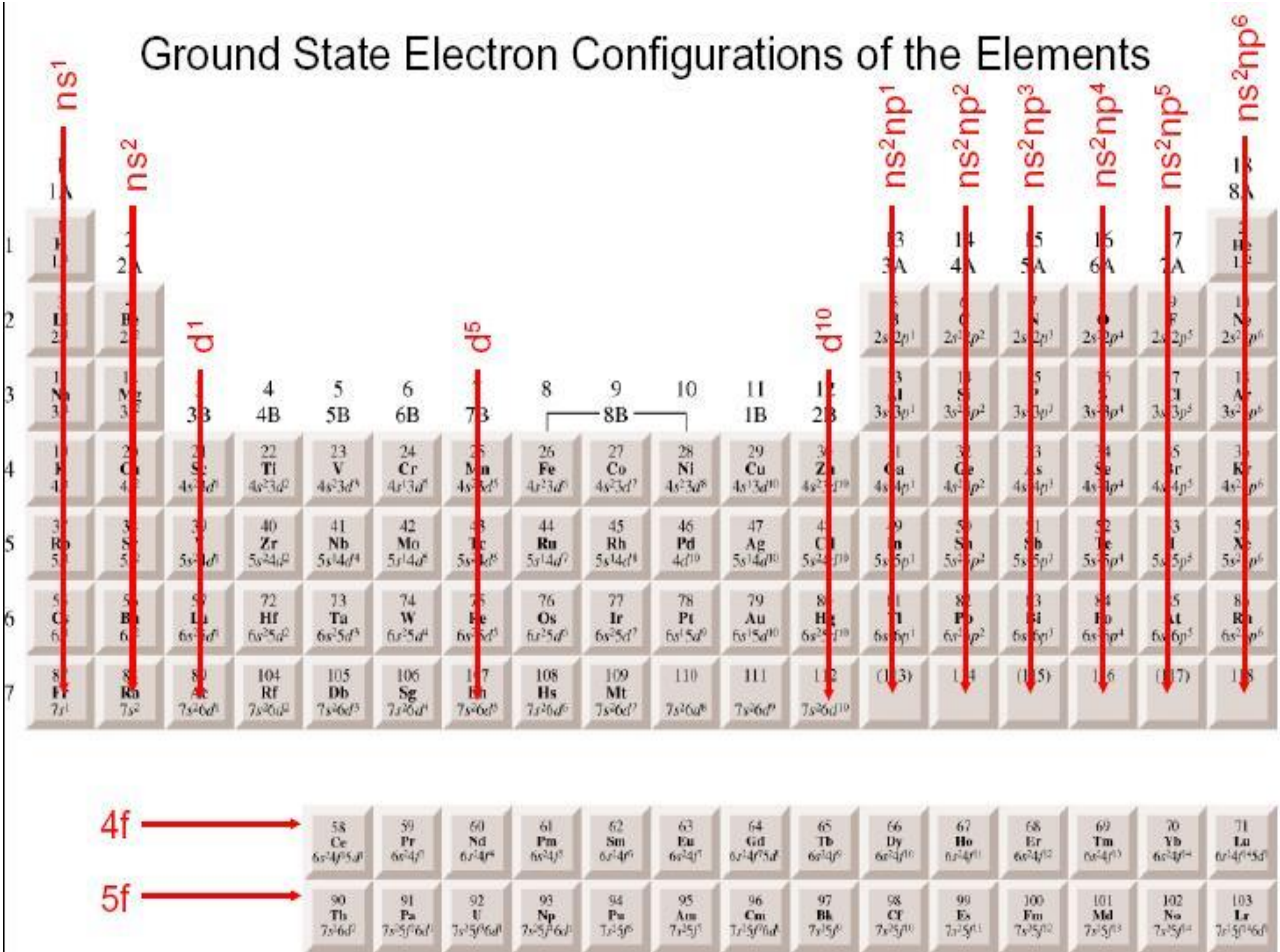
**(4) Inner –transition Elements (f-block Elements:**

The elements in which the extra electron enters the *(n-2)f-orbitals* of the *(n-2)* th main shell are called *f-block* element. The inner transition elements are all metals and show variable oxidation states. Their compounds are highly coloured. *f*-block elements are of two types:

**(a) *4f-series* (Lanthanides or Lanthanones).** This series has 14 elements: $Ce_{58}$ to $Lu_{71}$. In these 4*f*-orbital is being filled. These are called lanthanides or lanthanones.

**(b) *5f-series* (Actinides or Actinones).** This series also has 14 elements: $Th_{90}$ to $Lw_{103}$. These are called actinides or actinones. In these 5*f*-orbital is being filled.

# Ground State Electron Configurations of the Elements



# Variation of Properties within Periods and Groups
# (Periodic Trends)

1. **Variation of Metallic Character of the Elements:**

Chemical properties associated with metallic character result from how readily metals lose their electrons to form cations (positively charged ions).

Physical properties associated with metallic character include metallic luster, shiny appearance, high density, high thermal conductivity, and high electrical conductivity. Most metals are malleable and ductile and can be deformed without breaking. Many metals are hard and dense.

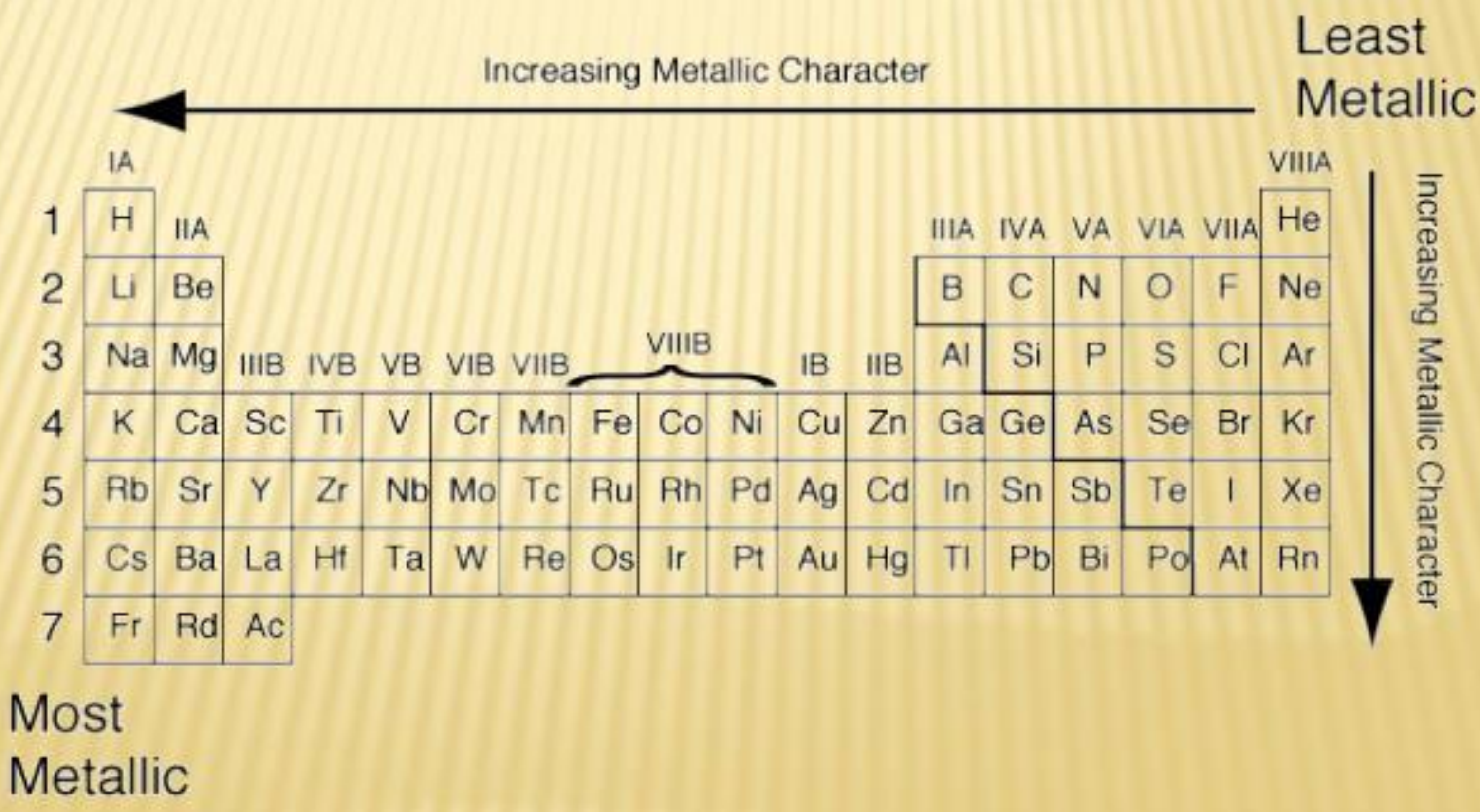
There are trends in metallic character as move across and down the periodic table. Metallic character decreases as move across a period in the periodic table from left to right. This occurs as atoms more readily accept electrons to fill a valence shell than lose them to remove the unfilled shell.

Metallic character increases as move down an element group in the periodic table. This is because electrons become easier to lose as the atomic radius increases, where there is less attraction between the nucleus and the valence electrons because of the increased distance between them.

For instance, the most non-metallic elements, fluorine, chlorine, oxygen, sulphur, nitrogen are found at the upper right of the Periodic Table whereas the most basic metals, the alkali and alkaline earth metals are at the lower left of the Table.

# PERIODIC TRENDS IN METALLIC CHARACTERISTICS

- Metallic character decreases across a period and increases down a group.



**2. Variation in the Atomic Size:**

Atoms do not have specific outer boundary; their **atomic size** is based on atomic radius, which is *one-half the Inter nuclear distance of two bonded identical atoms*. so when atomic radius is discussed as a periodic trend, what's usually meant is bonding atomic Radius.

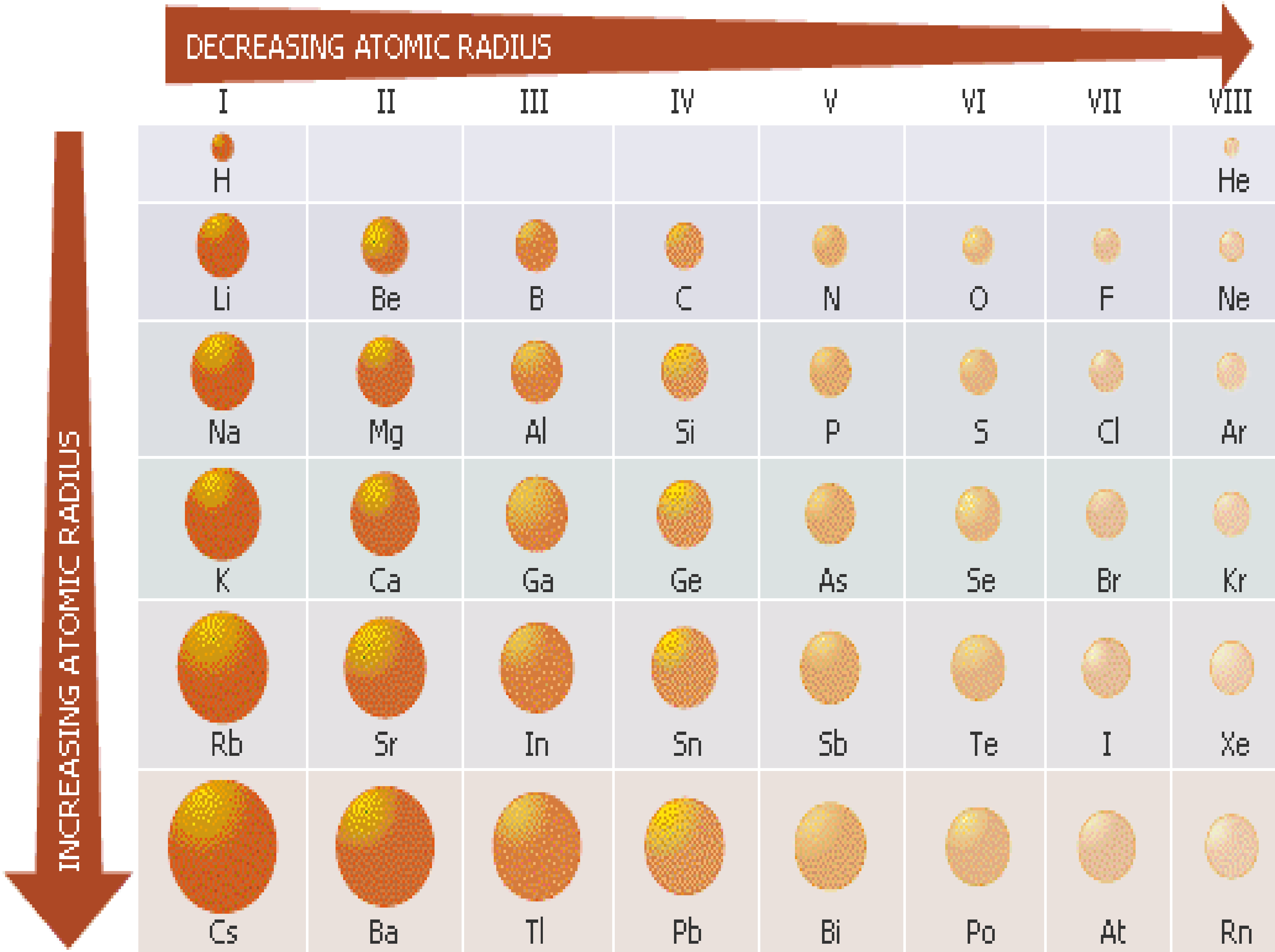
Two factors that greatly influenced atomic size are *principal quantum number* ($n$) of valence shell and the *effective nuclear charge* ($Z_{eff}$). Since principal quantum number $n$ increases down a group and $Z_{eff}$ increase significantly across a period, the following general trends in *atomic size* are observed:

Atomic radius generally increases from top to bottom down a group. This is a result of the increase in the principal quantum number, n, making the atom thus larger.

Atomic radius generally decreases from left to right across a period. This is mainly due to effective nuclear charge.

As electrons added to the same valence shell, also add protons to the nuclear charge. And thus across the Period nuclear charge predominates, and draws the valence electrons towards the nuclear core.

The greater the effective nuclear charge, the more the outer electrons feel that attraction from the nucleus, with the result of a marked decrease in atomic radius.


DECREASING ATOMIC RADIUS
INCREASING ATOMIC RADIUS
I
II
III
IV
V
VI
VII
VIII
H
He
Li
Be
B
C
N
O
F
Ne
Na
Mg
Al
Si
P
S
Cl
Ar
K
Ca
Ga
Ge
As
Se
Br
Kr
Rb
Sr
In
Sn
Sb
Te
I
Xe
Cs
Ba
Tl
Pb
Bi
Po
At
Rn

## 3. Variation in Ionic Radii:

Cations, or positively charged ions, are smaller than their "parent" atoms. This is because cations are formed when those outermost orbitals are vacated of electrons.

This also decreases electron-electron repulsions. Therefore, the resulting ions are smaller as there are not as many occupied orbitals and the effective nuclear charge affecting the remaining electrons increases, pulling electrons in more closely.

Anions, or negatively charged ions, are larger than their "parent" atoms. This is because electrons are added to form these ions, increasing electron-electron repulsions, making the electrons spread out more. Also, effective nuclear charge felt by the outermost electrons decreases.

The ionic sizes follow the same trends as the trends in atomic sizes for the neutral elements.

All the elements for each group of metals lose the same number of electrons, which means that the ionic sizes will be primarily affected by the number of energy levels in the electron cloud.

Since the number of energy levels still increases from top to bottom, the ionic size also increases down a group of elements in the periodic table.

For similar reasons, the trend across a period is the same for both ions and neutral atoms. All the metal elements in a given period will lose their outer shell electrons but still have the same number of core electrons.

As a result, the nuclear charge increases from left to right, while the number of core electrons remains the same. This means that the ion size will decrease from left to right across a period.

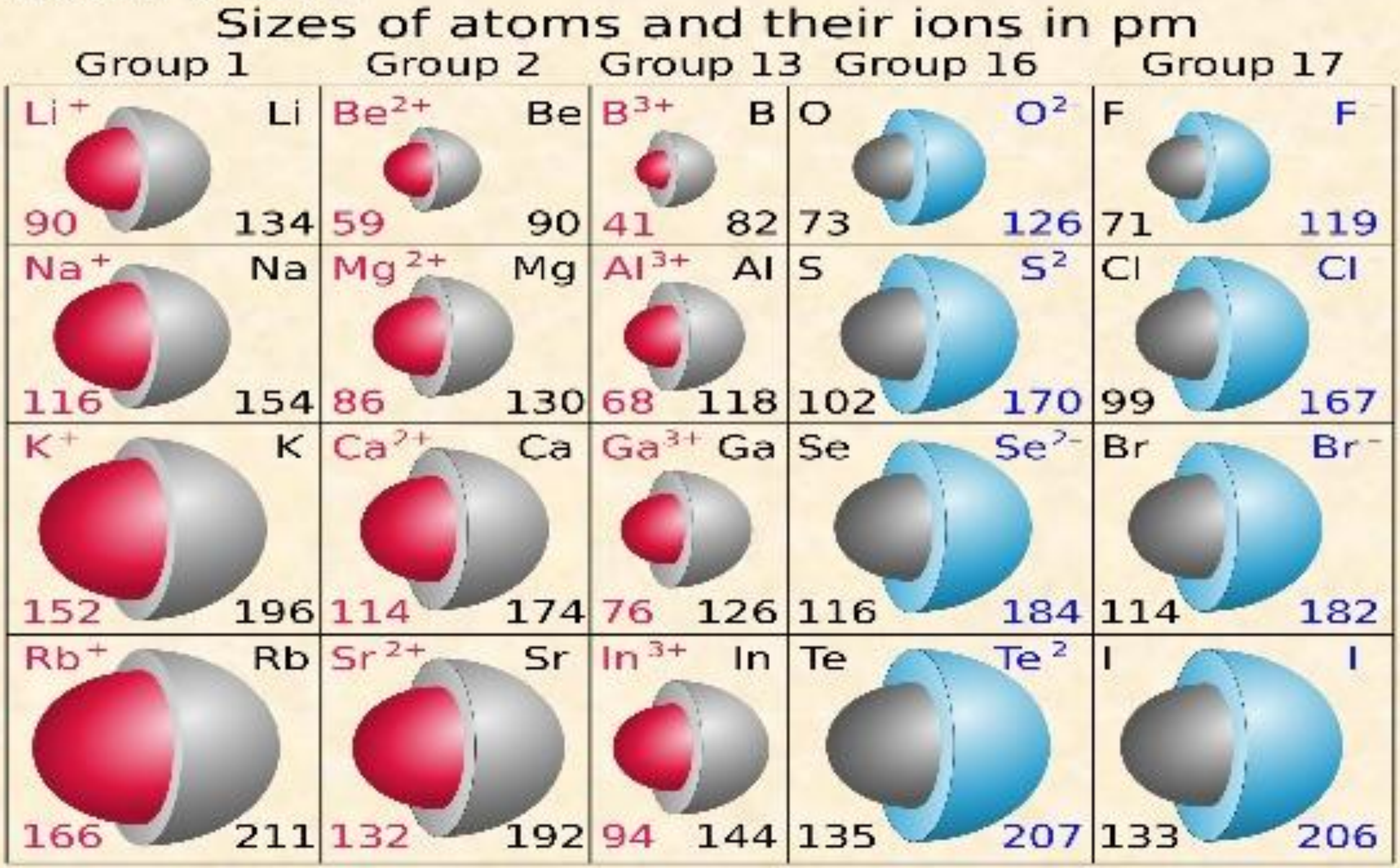
Non-metals also see the same trends in size as the neutral elements. The negative ions increase in size as move down a group and decrease in size as move from left to right across a period.

In other words, as go from top to bottom down a group ionic size increases or left to right across a period, the ionic size decreases *as long as* comparing all metals or all non-metals.

Between the metals and the non-metals, the ionic size increases because of switching from cations, which lose electrons, to anions, which gain electrons.

# Ionic Radius

**Ionic radius**, $r_{ion}$, is the **radius** of an atom's ion.



**4. Variation in the Ionization Energy:** The ionization energy is defined as the energy required to remove the outermost electron from an atom. Ionization energy of an atom is the most important fundamental property of an atom and the metallic character is largely determined by its ionization energy

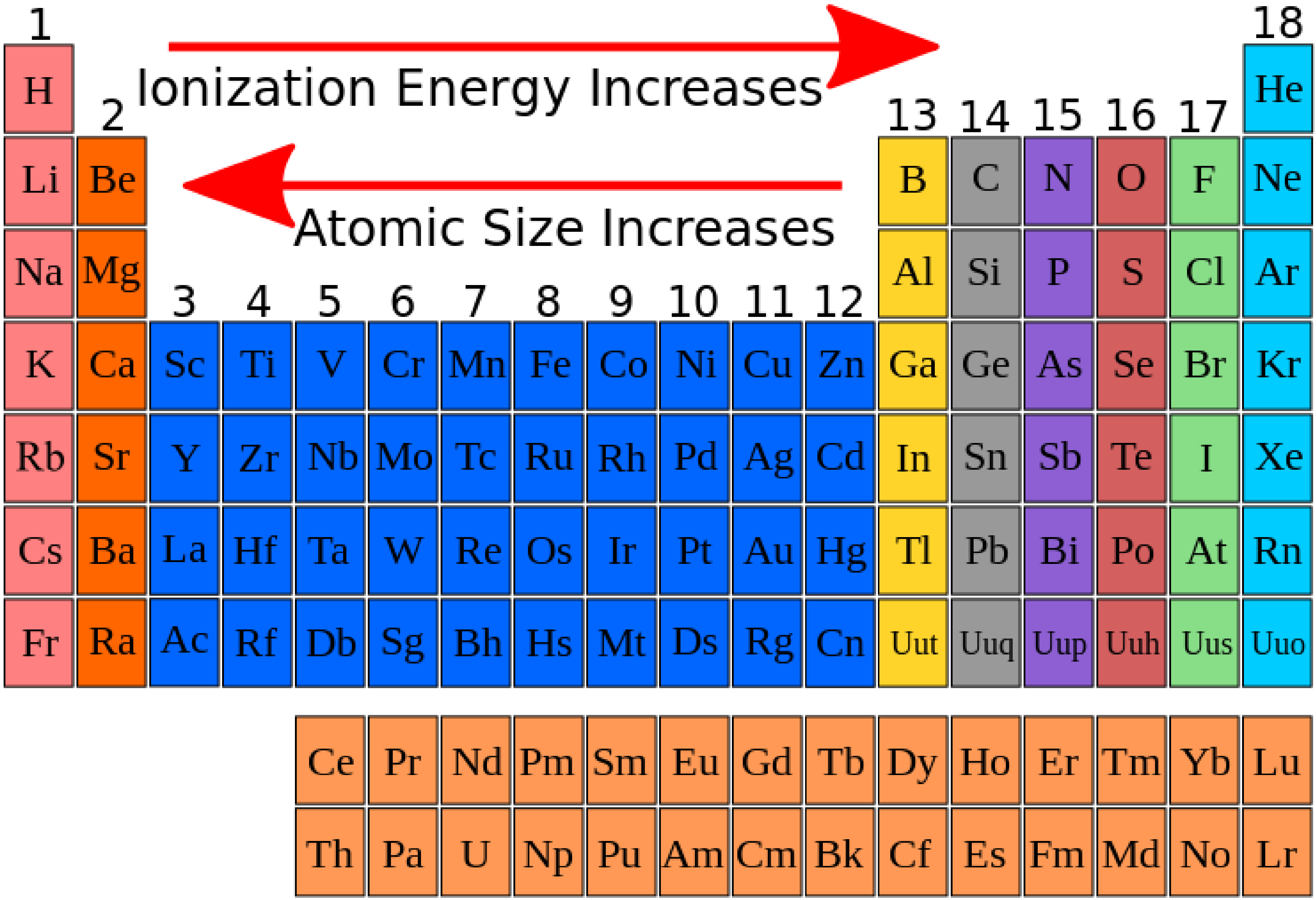
Since the IE of an elements is influenced by the $Z_{eff}$ of the atom, the following trends are observed in the periodic table:

IE generally *increases across a period*: $Z_{eff}$ increases and valence shell electrons are held more firmly.

IE *decreases down a group*: valence electrons are farther away from the nucleus and held less firmly.

The trends in IE are opposite those in atomic size: it is easier to remove an electron (lower IE) from a large atom than from a small one.

This indicates that metals give positive ions readily by the loss of electron and non-metals have no tendency to form positive ions during chemical reactions. Thus the ionization potential increases in a series and show decreasing tendency with a group in the periodic classification.



Ionization Energy Increases
Atomic Size Increases
1
18
2
13 14 15 16 17
3 4 5 6 7 8 9 10 11 12
H He
Li Be B C N O F Ne
Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba La Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Ac Rf Db Sg Bh Hs Mt Ds Rg Cn Uut Uuq Uup Uuh Uus Uuo
Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu
Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr

**5. Variation in Electron Affinities:**

Electron affinity, is a measure of the energy released when an electron is added to an atom to form a negative ion.

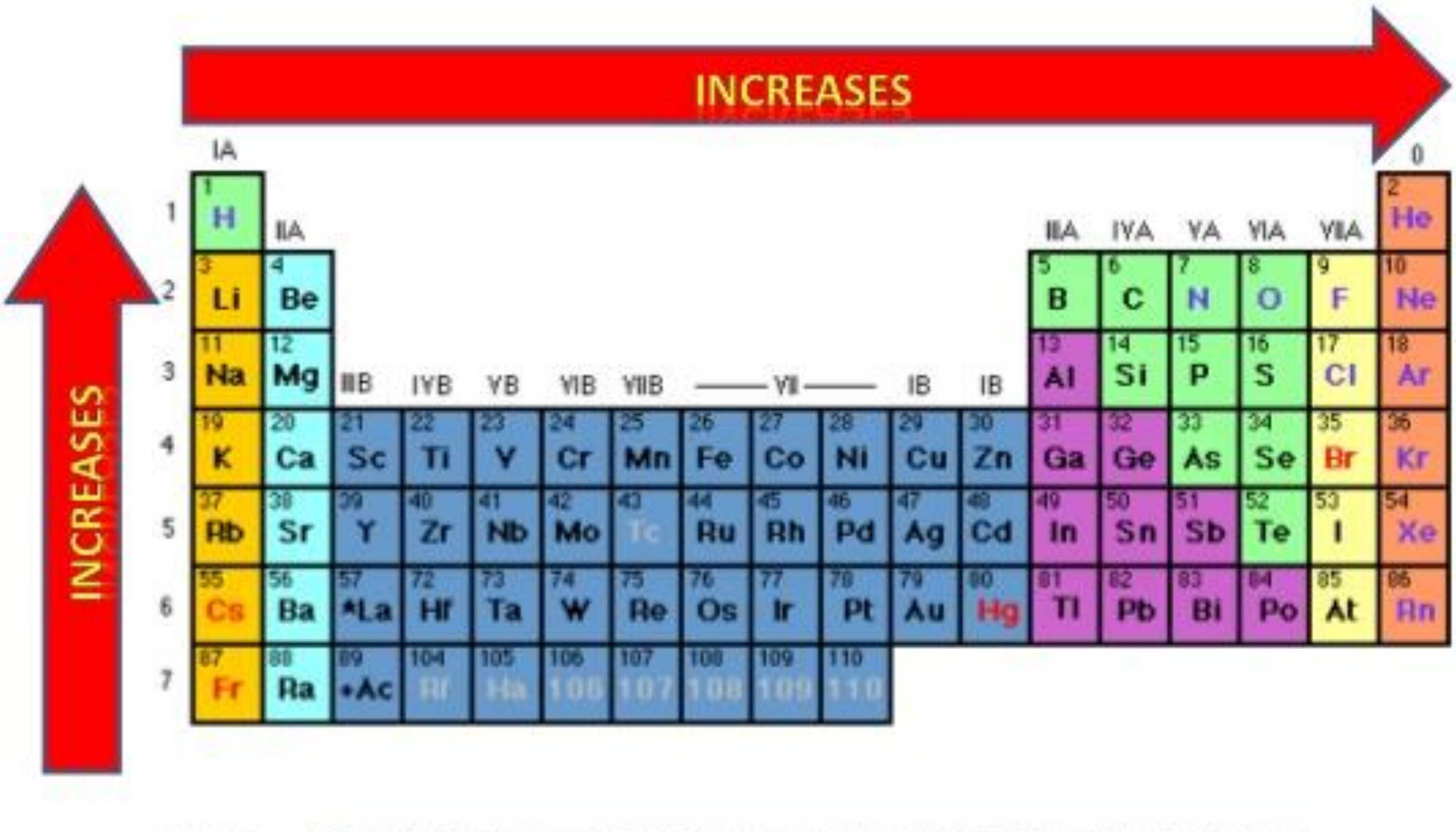
Metals obviously have small electron affinities and non-metals, on the other hand, have large values of electron affinities.

In general, electron affinity increases (or becomes *more negative*) from left to right across a period. This is due to increasing effective nuclear charge, which more readily pulls these new electrons in.

Electron affinity decreases (or becomes *less negative*) from top to bottom down a group. As we move down a group, the atomic radii increases, meaning these electrons are farther away from the nucleus and thus experience less of an electron-nucleus attraction.

# Periodic Trends: Electron Affinity

A measure of the change in energy that occurs when an electron is added to an atom.



**6. Variation in Electro-negativities:**
Electronegativity is a measure of the ability of an atom to attract the electrons when the atom is part of a compound.

Electronegativity differs from electron affinity because electron affinity is the actual energy released when an atom gains an electron. Electronegativity is not measured in energy units.

Since metals have few valence electrons, they tend to increase their stability by losing electrons to become cations. Consequently, the electro-negativities of metals are generally low.

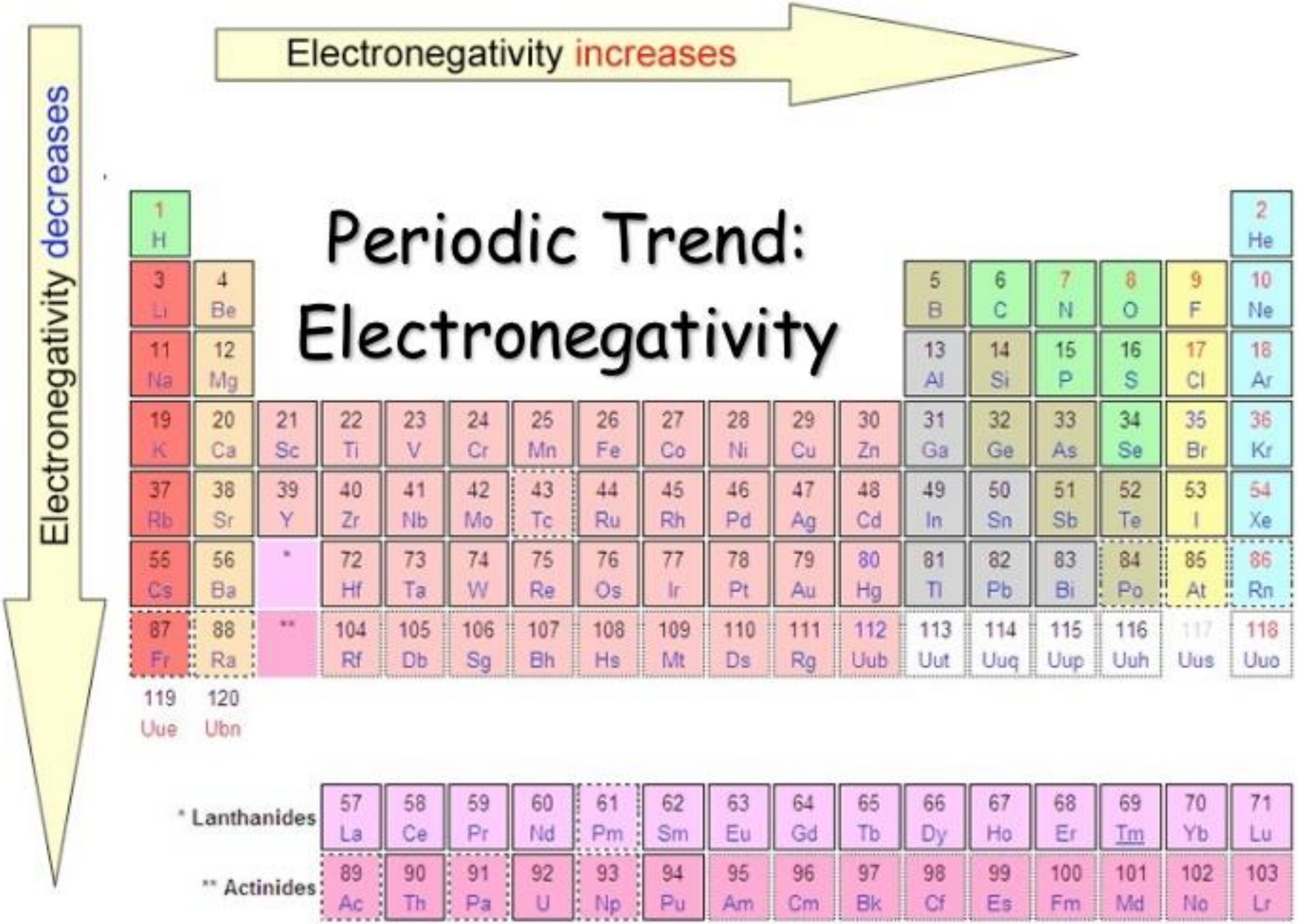
Non-metals have more valence electrons and increase their stability by gaining electrons to become anions. The electro-negativities of non-metals are generally high.

Electro-negativities generally increase from left to right across a period. This is due to an increase in nuclear charge.

Electro-negativities generally decrease from top to bottom within a group due to the larger atomic size.

Alkali metals have the lowest electro-negativities, while halogens have the highest.

Because most noble gases do not form compounds, they do not have electro-negativities.


Electronegativity increases
Electronegativity decreases
Periodic Trend:
Electronegativity
1 H
2 He
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
*
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
**
104 Rf
105 Db
106 Sg
107 Bh
108 Hs
109 Mt
110 Ds
111 Rg
112 Uub
113 Uut
114 Uuq
115 Uup
116 Uuh
117 Uus
118 Uuo
119 Uue
120 Ubn
* Lanthanides
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
** Actinides
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

**7. Variation in the Oxidising and Reducing Power:**

Oxidising substances have tendency to accept electrons and are converted into lower oxidation states.

The non-metals at the extreme right of the Periodic Table having high ionization potentials, electron affinities and electro-negativities tend to act as oxidising agents in chemical reactions with other substances.

Similarly, reducing substances give up electrons and are converted into higher oxidation state during chemical reaction.

The reducing power is the highest with the metals at the beginning of the periods, where the ionization potentials, electron affinities and electro-negativities are low. Thus, alkali metals have the greatest reducing power.

In general, the reducing power of the elements is progressively lower as we pass across the periods and higher, as we go down the groups. Francium should be the strongest reducing agent and fluorine, the strongest oxidising agent of all the elements.

# Properties of the transition metals

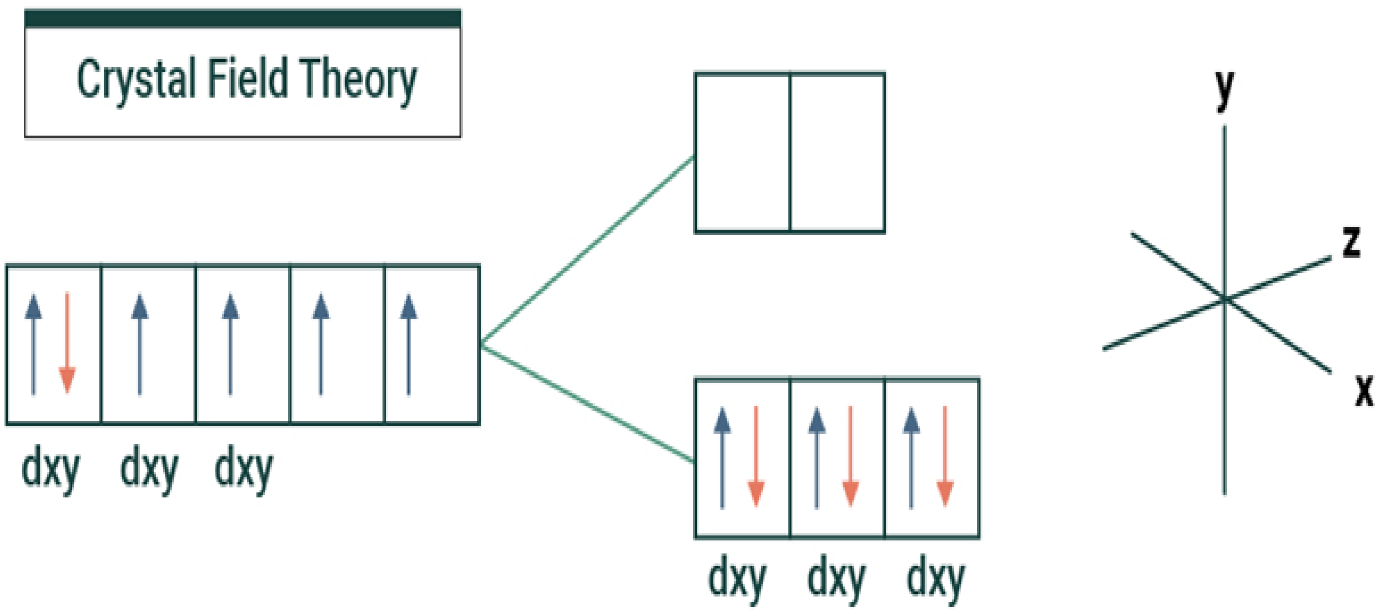
- These elements form coloured compounds and ions. This colour is explained by the d-d transition of electrons.
- There is a relatively low gap in energy between the possible oxidation states of these elements. The transition elements, therefore, exhibit many oxidation states.
- Many paramagnetic compounds are formed by these elements, because of the unpaired electrons in the d orbital.

- A large variety of ligands can bind themselves to these elements. Due to this, a wide variety of stable complexes are formed by transition elements.
- These elements have a large ratio of charge to the radius.
- Transition metals tend to be hard and they have relatively high densities when compared to other elements.

- The boiling points and the melting points of these elements are high, due to the participation of the delocalized d electrons in metallic bonding.
- This metallic bonding of the delocalized d electrons also causes the transition elements to be good conductors of electricity.

# Transition metals form coloured compounds

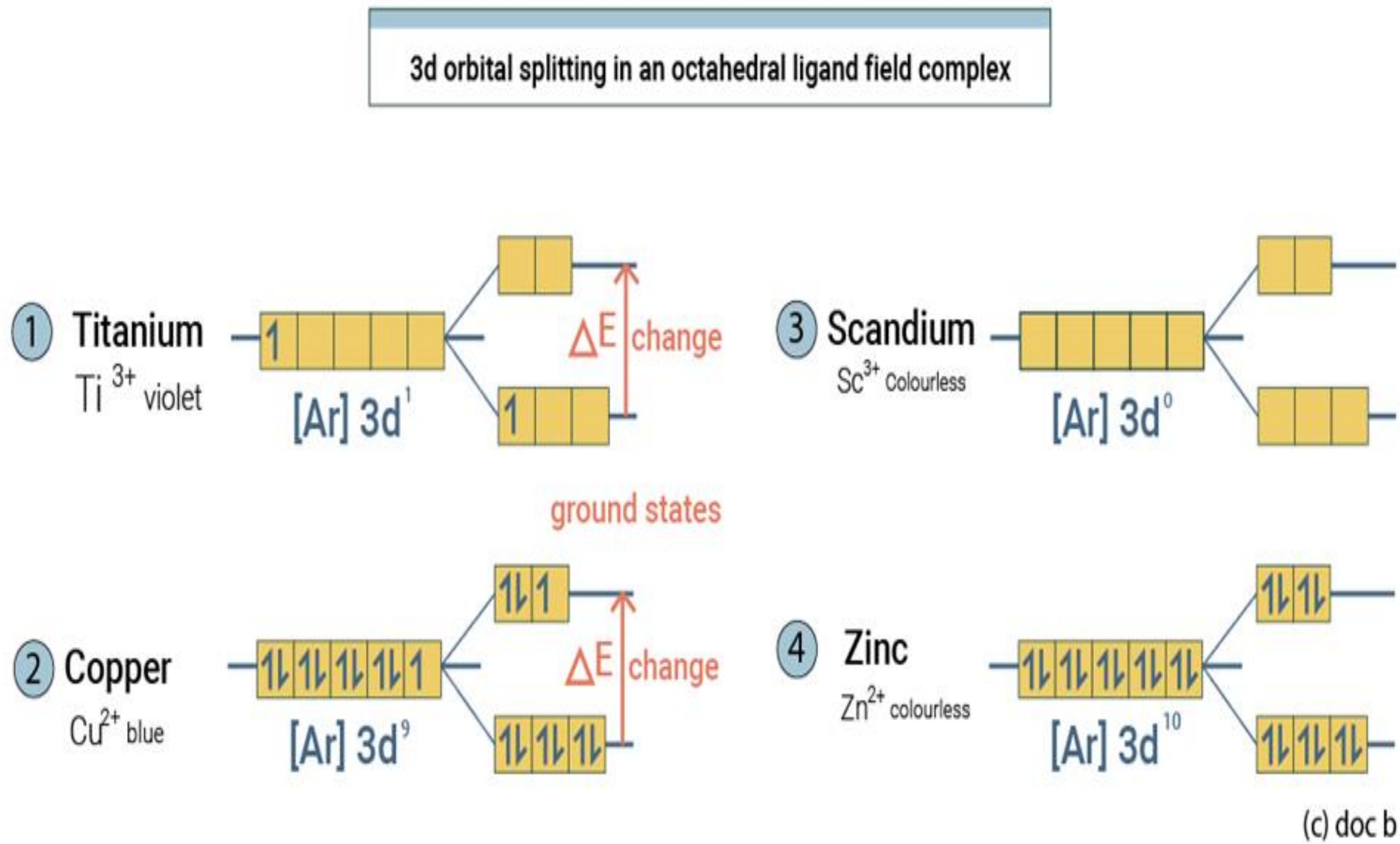
In transition metals the D block is easily split. This creates two small energy levels and as the D block elements are often unfilled this means there is space in the energy levels of the D block for electrons to be excited from the lower D block energy level to the higher energy level and when they come back down they emit photons of wavelengths in the visible region of the spectrum.




Degenerate ‘d’ orbitals split into two non-degenerate sets in an octahedral crystal (ligand) field

For example copper sulphate has a copper 2+ ion which has the electron configuration (Ar) $3d^9$ having 9 electrons in the d block. This means there is space in the split d orbitals for an electron to be excited into the upper d block energy levels when it is split and for it to then come back down to its original energy level emitting a photon.

Where as zinc whilst it is in the d block of metals it has a full d block. this means even though the d block splits into two energy levels as it has a full number of electrons, $3d^{10}$, this means no electrons can be promoted to a higher energy level as there is no space.


3d orbital splitting in an octahedral ligand field complex
1 Titanium
$Ti^{3+}$ violet
[Ar] $3d^1$
ΔE change
ground states
2 Copper
$Cu^{2+}$ blue
[Ar] $3d^9$
ΔE change
3 Scandium
$Sc^{3+}$ Colourless
[Ar] $3d^0$
4 Zinc
$Zn^{2+}$ colourless
[Ar] $3d^{10}$
(c) doc b

The reason the compounds is colored is when the electron that was excited into the higher energy level comes back down emits a photon. The wavelength of the photon will depend on the light which is absorbed.

The wavelength and frequency of the light that is emitted is effected by how big the energy gap is in the D block that has been split. The magnitude of this splitting of the d block orbitals is determined by the ligands, metal ion charge and the coordination number of the complex. The bigger the gap the higher the frequency of energy released as a photon.

# Noble gas

# Location and List of the Noble Gases on the Periodic Table

The noble gases, also known as the inert gases or rare gases, are located in Group VIII or International Union of Pure and Applied Chemistry (IUPAC) group 18 of the periodic table. This is the column of elements along the far right side of the periodic table. This group is a subset of the nonmetals.

The noble gases are:

- Helium (He)
- Neon (Ne)
- Argon (Ar)
- Krypton (Kr)
- Xenon (Xe)
- Radon (Rn)
- Oganesson (Og)

With the exception of oganesson, all of these elements are gases at ordinary temperature and pressure. There haven't been enough atoms produced of oganesson to know its phase for certain, but most scientists predict it will be a liquid or solid.

Both radon and oganesson consist only of radioactive isotopes.

## Noble Gas Properties

The noble gases are relatively nonreactive. In fact, they are the least reactive elements on the periodic table. This is because they have a complete valence shell. They have little tendency to gain or lose electrons. The noble gases have high ionization energies and negligible electronegativities. The noble gases have low boiling points and are all gases at room temperature.

## Summary of Common Properties

- Fairly nonreactive
- Complete outer electron or valence shell (oxidation number = 0)
- High ionization energies
- Very low electronegativities

- Low boiling points (all monatomic gases at room temperature)
- No color, odor, or flavor under ordinary conditions (but may form colored liquids and solids)
- Nonflammable
- At low pressure, they will conduct electricity and fluoresce

## Uses of the Noble Gases

The noble gases are used to form inert atmospheres, typically for arc welding, to protect specimens, and to deter chemical reactions. The elements are used in lamps, such as neon lights and krypton headlamps, and in lasers. Helium is used in balloons, for deep-sea diving air tanks, and to cool superconducting magnets.

Why noble gases are chemically inert?

When elements react, their atoms complete their outer shells by losing, gaining, or sharing electrons . The atoms of noble gases already have complete outer shells, they do not have free electrons to react or to make chemical bonds, so they have no tendency to lose, gain, or share electrons. This is why the noble gases are inert and do not take part in chemical reactions.

# Alkali metals

Alkali metals take up the leftmost side of the periodic table. The group 1 elements consist of elements:

- Lithium (Li)
- Sodium (Na)
- Potassium (K)
- Rubidium (Rb)
- Caesium (Cs)
- Francium (Fr)

They belong to the s-block elements of the periodic table as their outermost electron enters the s orbital giving them the electronic configuration of $ns^1$.

As the alkali metals have only 1 electron in their valence shell, they readily lose it, making them among the most reactive elements on earth. Thus, they are highly electropositive metals. They are called alkali metals because they form strongly alkaline hydroxides with water.

The alkali metals, found in group 1 of the periodic table (formerly known as group IA), are very reactive metals that do not occur freely in nature. These metals have only one electron in their outer shell. Therefore, they are ready to lose that one electron in ionic bonding with other elements.

Alkali metals are known for being some of the most reactive metals. This is due in part to their larger atomic radii and low ionization energies. They tend to donate their electrons in reactions and often have an oxidation state of +1. These metals are characterized as being extremely soft and silvery in color.

They also have low boiling and melting points and are less dense than most elements. Li, Na, and K have the ability to float on water because of their low density. All of these characteristics can be attributed to the large atomic radii and weak metallic bonding these elements possess.

Group 1 elements have a valence electron configuration is ns1 and are good reducing agents (meaning they are easily oxidized). All of the alkali metals are found naturally in nature, but not in their pure forms. Most combine with oxygen and silica to form minerals in the Earth and are readily mined as they are of relatively low densitys and thus do not sink.

# Limitation of Periodic Table

***1. Position of Hydrogen:*** The position of hydrogen in the periodic table is left undecided. It has similarities in properties with both the alkali metals and the halogens. According to the atomic number or atomic weight, hydrogen should occupy a position just before helium.

## Similarities to Metals

Hydrogen shares many similarities with alkali metals, i.e. elements in group I-A. This is one of the factors that dictates the position of hydrogen in the table. Let us take a look at the similarities

Electronic Configuration: Like all the elements of the group, Hydrogen also has one electron in its last shell, the valence shell. Let us take a look at the composition of valence shells of a few of these alkali metals.

H (z=1) : $K^1$

Li (z=3) : $K^2$, $L^1$

Na (z=11) : $K^2$, $L^8$, $M^1$

Good Reducing Agent: Hydrogen is a strong reducing agent like all the other alkali metals.

$Fe_2O_3 + 4\ H_2 \rightarrow 3Fe + 4H_2O$

$B_2O_3 + 6\ K \rightarrow 2B + 3\ K_2O$

Forms Halides: Also just like alkali metals, hydrogen combines with electronegative elements to form halides

$2Na + Cl_2 \rightarrow 2NaCl$

$H_2 + Cl_2 \rightarrow 2HCl$

## Differences with Metals

Non-metal: Hydrogen is essentially not a metal like all alkali metals, but a non-metal

Loss of Electron: Although it has only one electron in its outer shell, hydrogen cannot easily lose this electron to gain electropositivity. All other alkali metals can do this with ease.

State: At room temperatures where all alkali metals exist is the solid state, hydrogen is a gas.

Size of Atom: The H+ ion of hydrogen is much smaller than ions of alkali metals.

Ionization Potential: The ionization potential of hydrogen is over 300 Kcal per mole, The maximum ionization potential for metals is 147 Kcal per mole.

Similarities to Halogens

Noble Gas Configuration: Hydrogen can gain one electron to complete its valence shells. Halogens also have seven electrons in their last shell and can gain one electron to gain noble gas configuration.

Electronegativity: They also share the same electronegative nature. Hydrogen also gains one electron (not looses) to become stable and so do halogens.

$H + e^{-} \rightarrow H^{-}$

$Cl + e^{-} \rightarrow Cl^{-}$

Diatomic Molecules: Both hydrogen and halogens form diatomic molecules. Hydrogen forms $H_2$ and, halogens are $Cl_2$, $F_2$ etc

Reaction with Metals: Hydrogen combines with metals to form metallic hydrides. Similarly, halogens also combine with metals to form metal halides.

$2Na + H_2 \rightarrow 2NaH$

$Ca + H_2 \rightarrow CaH_2$

Covalent Bonding: Halogens and hydrogen both also combine with non-metals to form molecules with covalent bonding.

## Differences with Halogens

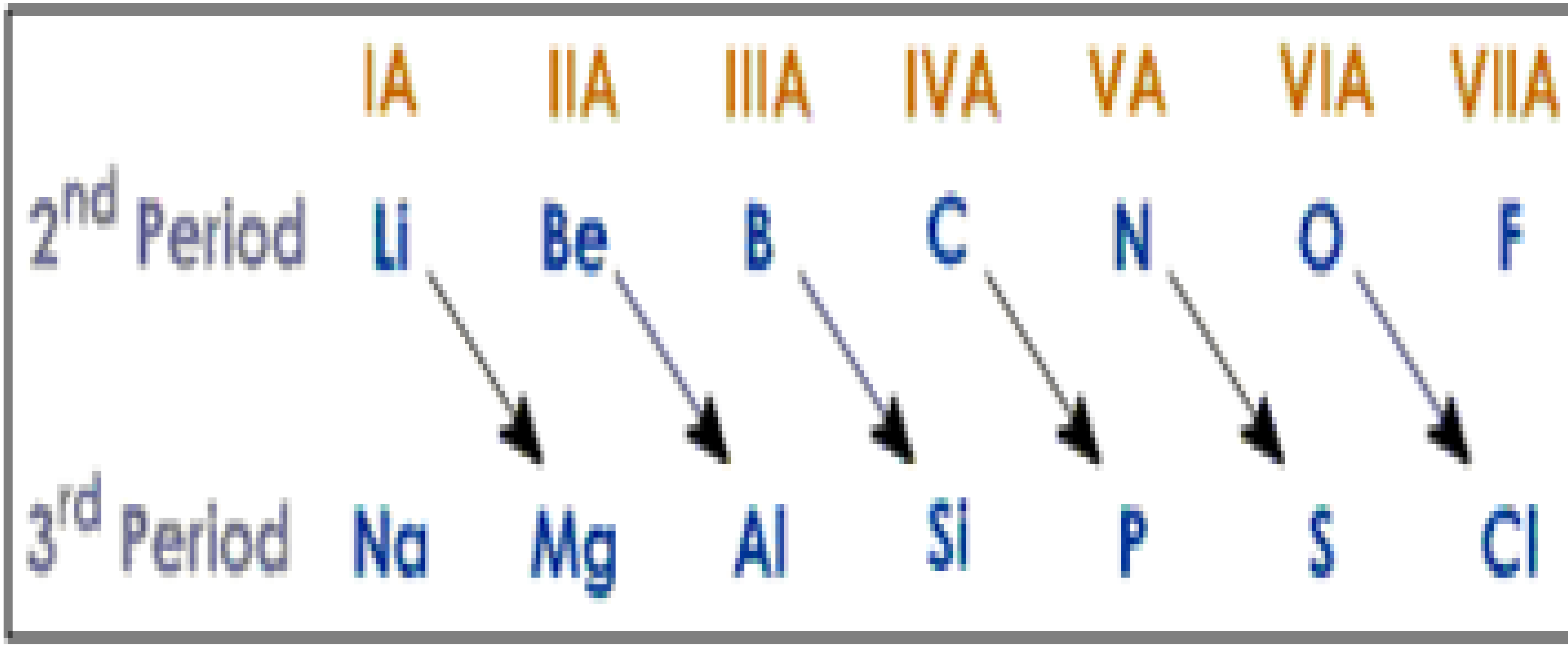
Structure of Atom: Hydrogen has only one electron in its outer shell. All halogens have seven electrons in their last shell

Size of Atom: The size of the H- ion is much larger than those of the ions of Halogens. This is because hydrogen has only one electron and proton and the pull of the nucleus is less.

Reaction with Water: Also unlike halogens, the hydrogen ion H- is unstable in water.

However, the properties of hydrogen do not completely match any of these groups. Due to this reason, the position of hydrogen is still considered as undecided as hydrogen behaves like halogens as well as like alkali metals.

***2. Diagonal Relationships:*** The similarity in properties observed between two elements placed diagonally from left to right in two adjacent periods and two adjacent groups in the periodic table is known as the diagonal relationship.




Diagonal relationship on the Periodic Table

The diagonal relationship exists between the certain pairs of diagonally adjacent elements in the second and third periods of the periodic table.

The chemistry of lithium is similar to the chemistry of magnesium. Chemistry of beryllium is similar to that of aluminum and the chemistry of boron is similar to the chemistry of silicon.

| | Li | Be | B | C |
|---|---|---|---|---|
| Charge of ion | +1 | +2 | +3 | +4 |
| Ionic radius, Å[a] | 0.73 | 0.41 | 0.25 | 0.29 |
| Charge density | 1.4 | 4.9 | 12 | 14 |
| Electronegativity | 1.0 | 1.5 | 2.0 | 2.5 |

| | Na | Mg | Al | Si |
|---|---|---|---|---|
| Charge of ion | +1 | +2 | +3 | +4 |
| Ionic radius, Å | 1.13 | 0.71 | 0.53 | 0.40 |
| Charge density | 0.88 | 2.8 | 5.7 | 10 |
| Electronegativity | 0.9 | 1.2 | 1.5 | 1.8 |

On crossing the period of the periodic table, the size of the atom is decreased and on descending the group in periodic- table size of the atoms is increased.

The other explanation is based on the sizes of the ions formed by the removal of valence electrons. Thus, **$Li^{+}$** ion is almost of the same size as **$Mg^{+2}$** ion. Similarly, **$Be^{+2}$** and **$Al^{+3}$** ions have approximately the same ionic size. **$B^{+3}$** and **$Si^{+4}$** also present the same situation. Compounds having similar properties of the elements showing diagonal relationship are formed due to the effect of ionic sizes.

Moving along the periods the elements progressively becomes more covalent, more electronegative and less reducing.

Whereas on descending the group in the periodic table the elements become more basic, more ionic and less electronegative.

Thus, **Li** in group **IA** is more electropositive than **Be** in group **IIA**, but **Mg** is also more electropositive than **Be**. Thus, both **Li** and **Mg** are more electropositive than **Be** and less electropositive than **Na**.

So, both crossing and descending a group by one element cancel the changes and elements with similar chemistry and similar properties are often found.

However, after boron and silicon, it is not noticeable. Still, the reasons for diagonal relationship are not fully understood.