-->notes by AVISHAKE Sir

PERIODIC TABLE

Older methods of classification:

State dobereiners law of triads .Give one example of a triad.

Dobereiner Law of Triads, Dobereiner grouped the elements in triads (groups of three elements), in such a way that the middle element of the triad had both atomic mass and properties roughly equal to the average of the other two elements of the triad.

Examples :

(i) C (12) N (14) and O (16) : $\frac{12+16}{2} = 14$ |

(ii) Li(7), Na(23), K(39) : $\frac{7+39}{2} = 23$

Drawbacks :However, this method was soon discarded, since only a few elements known at that time could be arranged in such triads.

What is Newlands Law of Octaves.?

Newland law of Octaves: Newland gave the law of octaves, which states that : 'When elements are arranged in an increasing order of their atomic mass, every eighth element beginning from any element resembles the first element in its physical and chemical properties.

This method was also discarded, sine it failed to accommodate the heavier elements.

Mendeleevs Periodic Table:

Mendeleev published a table of elements called it Mendeleev's periodic table . He arranged the elements in the increasing order of their atomic masses. This arrangement enabled Mendeleev to place elements in vertical columns known as groups and in horizontal rows known as periods.

STATE MENDELEEFF'S PERIODIC LAW:

The physical and chemical properties of elements are - periodic functions of their - atomic weights.

Series	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
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	

Mendeleev's Original Periodic Table of Elements - [first three periods shown above]

- **MENTION MENDELEEV'S CONTRIBUTIONS .**

  - Elements were arranged - In increasing order of atomic weights - in horizontal rows called 'periods' and vertical columns called 'groups'.
  - Elements which are similar with respect to their chemical properties – are grouped together and have atomic weights of nearly the same value.
  - Elements in the same group - had the same 'valency' and similar chemical properties.
  - Based on the periodicity of properties - a number of gaps were left in the table - for undiscovered elements i.e. elements now discovered e.g. Scandium, Gallium and - Germanium originally called eka-boron, eka-aluminium and eka-silicon respectively.
  - The properties of the undiscovered elements - left in the vacant gaps was predicted.
  - Incorrect atomic weights of some of the arranged elements - were corrected with the knowledge of the atomic weights of the adjacent elements.

MENTION SOME OF THE DEFECTS OF MENDELEEFF'S PERIODIC TABLE

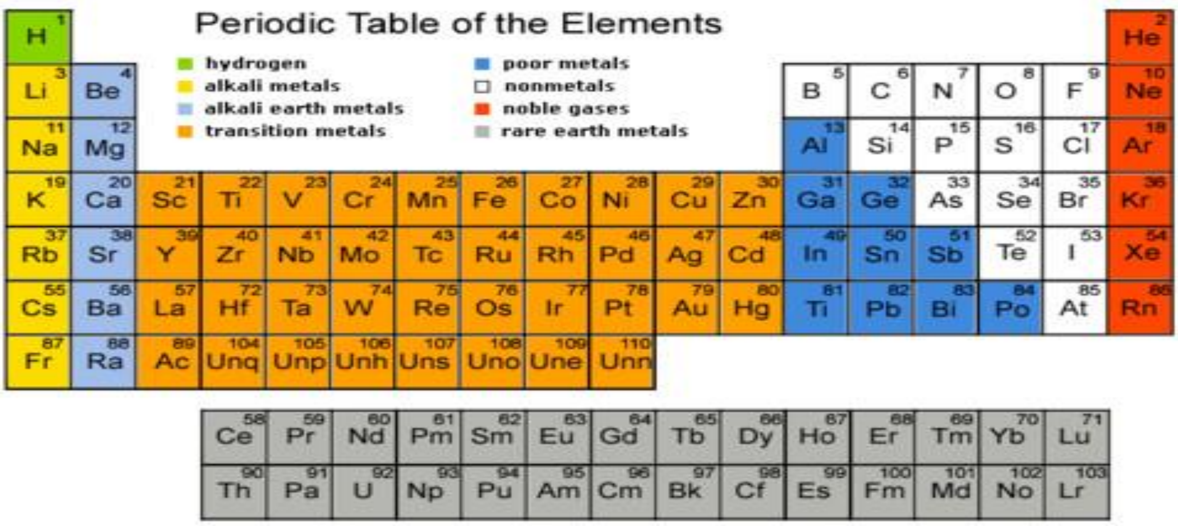
- Certain pairs of elements having higher atomic weights have been given positions before the elements having lower atomic weights.
- This defect disappears if elements were arranged according to their - atomic numbers. e.g. Co [at. wt. 58.9, at. no. 27) was placed before Ni [at. wt. 58.6, at. no. 28).
- Position of rare earths & actinides - Could be justified only if arranged according to the its - atomic numbers.
- Position of isotopes - Isotopes had to be placed in same position according to - atomic numbers

<u>Modern Periodic Table</u>

The defects of Mendeleev periodic table were removed by Henry Moseley who put forward the **<u>modern periodic law in 1913,</u>**

<u>State the modern periodic law:</u>

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



MENTION THE SALIENT FEATURES OF THE MODERN PERIODIC TABLE.

What are groups and periods .

Groups: The vertical columns in the periodic table are called as groups

Periods: The horizontal rows in the periodic table are called as periods.

GROUPS:

(1) The modern periodic table has eighteen vertical columns. They are known as groups, arranged front left to right in the order : IA. IIA, IIIB. IVB, VB, VIB. VIIB, VIII (three columns), IB, IIB. IIIA. IV-A. VA, VIA, VIIA and Zero.

Note : According to the recommendation of International Union of Pure and Applied Chemistr (1UPAC), the groups are numbered from I to 18.

OLD NOTATION	NEW NOTATION
IA	1
IIA	2
IIIB	3
IVB	4
VB	5
VIB	6
VIIB	7
VIII	8 9 10
IB	11
IIB	12
IIIA	13
IVA	14
VA	15
VIA	16
VIIA	17
0	18

A group is determined by the number of electrons present in the outermost shell. For example : Sodium has atomic number 11 and its electronic configuration is 2, 8, **1. It has one electron in the outermost orbit hence it is placed in group 1**. Similarly, magnesium, atomic number 12, electronic configuration 2, 8, 2 is placed in group 2

Group 1 : Alkali metals **-** They form strong alkalis with water.
Group 2 : Alkaline earth metals **-** They form weaker alkalis as compared to group 1.
Group 13 : Boron family **-** Boron is the first member of the group.
Group 14 : Carbon family **-** Carbon is the first member of the group.
Group 15 : Nitrogen family.
Group 16 : Oxygen family
Group 17 : Halogen family **-** The elements of this group form salts.
Group 18 :Inert gases –The elements of this group are inert and unreactive.

Periods

There are seven horizontal rows in the modern periodic table. They are known as periods. The number of shells present in an atom determines its period. For example

Elements of **period one** have **one shell** ,elements of **period two** have **two shells**, and that of **period three** have **three shells** and so on.

(1) The first period contains only two elements, (atomic nos. 1 and 2). It is the shortest period.

(2) The second and the third periods contain eight elements each (atomic nos. 3-10 in the second and atomic nos. 11-18 in the third period). These are short periods.

Diagonal relationship : The elements of the second period show resemblance in properties of the elements of the next group of the third period, leading to a diagonal relationship, viz. Li & Mg, Be & Al, B & St. These elements are called bridge elements.

Group →	1	2	13	14
Period 2	Li	Be	B	C
Period 3	Na	Mg	Al	Si

Typical elements

The third period elements, Na, Mg, Al, Si, P, S and Cl, summarise the properties of their respective groups and are called typical elements.

(3) The fourth and the fifth periods contain eighteen elements each (atomic nos. 19-36 in the fourth and atomic nos. 37-54 in the fifth period). These are long periods.
(4) The sixth period contains 32 elements (atomic nos. 55-86). It is the longest period.
(5) The seventh period (atomic nos. 87 onward) is yet an incomplete period.
(6) In Group 3 of the sixth period, there is a set of elements with atomic numbers 57 to 71 (La - Lu), beginning with lanthanum (La-57). They are known as lanthanides (rare earth elements).
(7) In Group 3 of the seventh period, there is a set of elements with atomic numbers 89 to 103 (Ac - Lr), beginning with actinium (Ac-89). They are known as actinides (radioactive elements).

Lanthanides and actinides have similar properties because they belong to the same Group 3. They are shown at the bottom of the periodic table because they are large in number, and to show them in the main body of the table will distort its shape.

TYPES OF ELEMENTS

1)What are Representative Elements
Elements of group 1, 2, 13, 14, 15, 16, 17 and 18 are called as representative elements.

2)What are Transition Elements. Give characteristics
Elements of group 3, 4, 5, 6, 7, 8, 9, 10, 11, 12 are called as transition elements.

Characteristics of transition elements :
They are metals with high melting and boiling point
Good conductors of heat and electricity
Most of these elements are used as catalyst
Most of these elements exhibit variable valency.
Many of the transition metals form coloured compounds.
Last two shells are Incomplete.

3)What are Inner Transition Elements.Give characteristics.
These elements are in the 6^{th} and 7^{th} period in group 3. They are called as lanthanides in period 6 and actinides in period 7.
Characteristics ::
Heavy metals with variable valencies.
All actinides are radioactive in nature
They also form coloured ions.
Last three shells are incomplete.

4)What are Inert Gases?
Elements of group Zero or group 18 are the inert gases which have a stable electronic configuration of 2 for helium and 8 in the last shell for the other gases. They do not react with the other elements. All of them are gases

<u>What is PERIODICITY?</u>
The properties that reappear at regular intervals, or in which there is gradual variation (i.e. increase or decrease) at regular intervals, are called 'periodic properties' and the phenomenon is known as the periodicity of elements.

<u>What are the Cause of periodicity?</u>
The cause of periodicity is the recurrence of similar electronic configuration i.e. same number of electrons in the outermost orbit.
In a particular group, electrons in the outermost orbit remain the same i.e. electronic configuration is similar. Since chemical properties of elements depend upon the number of electrons in their outermost shell, thus elements of the same group have similar properties.
For example in **<u>group 17, i.e., halogens</u>**, all elements have seven electrons (see electronic configuration given in 1.4) in their respective outermost shells. therefore. they show similar properties, such as:

<u>SHELLS (ORBITS) AND VALENCY</u>

Orbits : Electrons revolve around the nucleus in certain definite circular paths called orbits or shells.

(1) Number of shells

(a) Down a group, i.e., from top to bottom.
The number of shells increases successively, i.e., one by one, such that the number of shells that an element has, equals the number of the period to which that element belong

For example, in halogens (Group 17).

Element with atomic numbers	No. of shells	Electronic configuration K L M N O P	Period to which the element belongs
F (9)	2	2, 7	Second
Cl (17)	3	2, 8, 7	Third
Br (35)	4	2, 8, 18, 7	Fourth
1(53)	5	2, 8, 18, 18, 7	Fifth
At (85)	6	2, 8, 18, 32, 18,7	Sixth

(b) **Across a period**
On moving from left to right in a given period, the number of shells remains the same. For example, in the 2nd period, the number of shells remains two. i.e., equal to the number of the period.

<u>Valency</u>
Valency denotes the combining capacity of an element is equal to the number of electrons an atom can donate or accept or share.

On moving down a given group the number of electrons in the outermost shell are the same, therefore **valency in a group is the same**.

In a given period the number of electrons in the valence shell increase from left to right. The valency increases up to group 14 where it becomes 4 then it decreases to become 1 in group 17.

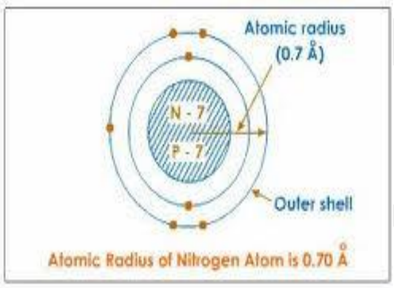
Elements of the 2^{nd} period	Li	Be	B	C	N	O	F	Ne
Atomic No.	3	4	5	6	7	8	9	10
Valency	1	2	3	4	-3	-2	-1	0

PERIODIC PROPERTIES

According to the modern periodic law, the properties of elements depend upon the following features
Atomic size (atomic radius)

(i) Metallic character
(ii) Non-metallic character
(iii) Ionization energy (ionization potential)
(iv) Electron affinity
(v) Electronegativity.

I) Atomic size (atomic radius):



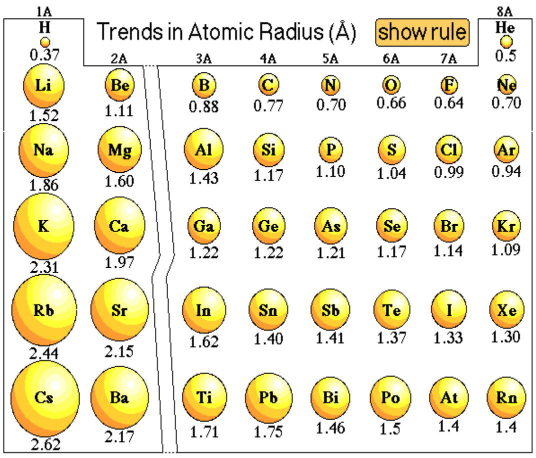
It is the distance between the centre of the nucleus of an atom and its outermost shell.
Atomic size depends upon:

(i) Number of shells and ii) nuclear charge

Unit : Angstrom : 1A = 10^{-10} m
Picometre : 1 pm = 10^{-12} m

(i) **Number of shells**
An increase in the number of shells increases the size of the atom because, the distance between the outermost shell and the nucleus increases.

(ii) **Nuclear charge**
It is the positive charge present in the nucleus of an atom, which is equal to the number of protons in the nucleus, i.e. the atomic number.
Any increase in nuclear charge decreases the size of the atom because the electrons are then attracted towards the nucleus with a greater force, thereby bringing the outermost shell closer to the nucleus.



Variation in atomic size (atomic radius)

(a) **In a group**.

In a group, the size of an atom increases as one proceeds from top to bottom. This is due to the successive addition of shells as one moves from one period to the next.

(b) In a period.

In a period, the size of an atom decreases from left to right. This is because the nuclear charge, i.e. the atomic number increases from left to **right in the same shell**, thereby bringing the outermost shell closer to the nucleus.

Note:

However the size of the atom of an inert gas is bigger than that of the preceding halogen atom. Therefore, considering the second period it has been found that, Lithium (Li) has the largest atomic size while Fluorine (F) has the smallest.

Neon has its outermost shell completely filled (inert gas) i.e. it has structural stability of its outermost shell consisting of an octet of electrons.

Why is a cation smaller then the parent atom.

Reason:

Cation is formed by the loss of electrons hence protons are more than electrons in cation, so electrons are strongly attracted by the nucleus and are pulled inward, so size decreases.

$Na - e^- \longrightarrow Na^+$

(2, 8, 1) (2, 8)

Why is an anion larger than the parent atom.

Reason

Anion is formed by gain of electron (s). the effective positive charge in the nucleus is less, so less inward pull is experienced hence size expands

$Cl + e- \longrightarrow Cl^-$

(2, 8, 7) (2, 8, 8)

Explain Metallic character:

Those elements which have a tendency to lose their valence electrons (electrons of the outermost orbit) and form a positive ion, are considered as metals.

$Na - e^- \longrightarrow Na^+$

(2, 8, 1) (2, 8)

The metallic character of elements depends on:

(i) atomic size

(ii) nuclear charge

(a) **Atomic size** : the greater the atomic size, the father the outermost orbit, and thus lesser the nuclear pull exerted on the electrons by the nucleus and hence electrons can be removed easily which makes it more metallic.

(b) **Nuclear charge** : greater the nuclear charge, greater is the force exerted by the nucleus on the electrons of the outermost orbit.

How does Metallic character vary down the group and across a period.?

On moving down a group the atomic size increases, the nuclear charge also increases. The effect of

an increased atomic size is greater as compared to the increase nuclear charge therefore metallic character increases as one moves down a group
Example: Lithium is the least metallic element and Francium is most metallic element in group I.

Across a period, atomic size reduces, hence elements cannot lose electrons easily therefore the metallic nature decreases across a period.
Example :
In the second period Lithium is the most metallic.
In the 3rd period Sodium is the most metallic.

Explain Non-Metallic character
Those elements which have tendency to gain electrons in their outermost orbit are considered as non-metals.
Non-metals usually have 5, 6 and 7 electrons in their outermost shells.
Non-metallic character depends on :

a) atomic size
b) nuclear charge

a) Atomic size
Smaller the atomic size, greater is the nuclear pull, tendency to gain electrons increases (non-metallic character increases).

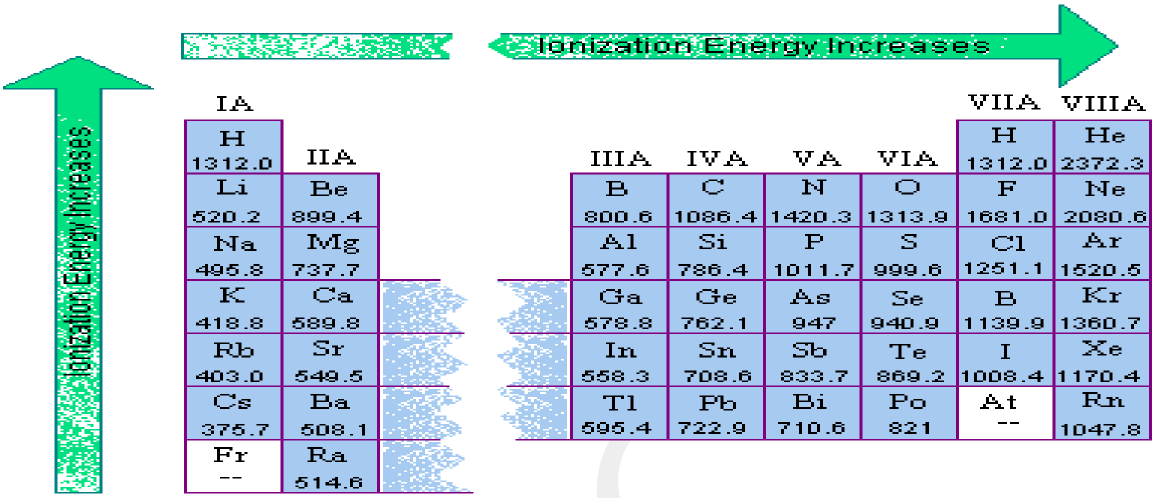
b) Nuclear charge
Greater the nuclear charge, greater is the tendency to attract electrons, hence more non-metallic is the element.
How does Non -metallic character vary down the group and across a period.?

Down a group: atomic size increases due to new shells added. The non-metallic nature decreases.
Across a period: non-metallic character increases since the atomic size decreases

Explain Ionisation energy or **Ionisation potential** or **Ionisation enthalpy**

The energy required to remove an electron from a neutral isolated gaseous atom to convert it to a positively charged gaseous ion is called ionization energy (I.E.) unit is measured in electron volts per atom and its SI unit is kJ per mole.
Note, the energy required to remove the 2nd electron is called the 2nd ionization energy and so on.
The ionization energy increases as the charge on the cation increases gradually because the nuclear charge is distributed over lesser no of electrons.
$IE1 < IE2 < IE3$ . . . . . .

Ionisation energy depends
Atomic Size : Greater the atomic size, lesser is the force of attraction, hence ionization energy is less.
Nuclear charge : Greater the nuclear charge, greater is the attraction of the electrons to the nucleus, hence the ionization energy increases.



How does Ionisation energy vary across the period and down the group?

Across a period: Ionisation energy tends to increase as one moves from left to right because the atomic size decreases due to the increase in the nuclear charge and thus more energy is required to remove the electrons.
<u>Down a group</u>: There is an increase in the atomic number (nuclear charge), the atomic size increases because of the addition of extra shells. This increase in the atomic size overcomes the effect of an increase in the nuclear charge, therefore ionization energy decreases with an increase in the atomic size.

Note : Helium has the highest ionization energy while Caesium will have the lowest ionization energy since Francium is a radioactive element.

<u>Explain Electron affinity or electron gain enthalpy</u>.
The amount of energy released when converting a neutral gaseous isolated atom to a negatively charged gaseous ion by the addition of electrons is called electron affinity.

$X_{(g)} + e^- \longrightarrow X^-_{(g)} + E.A.$

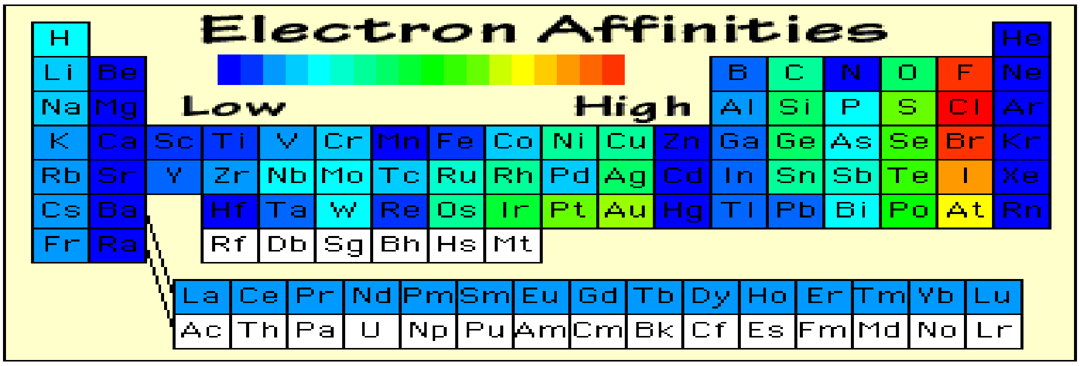
X is any element taken in its gaseous state.
Units are electron volts per atom (eV/atom) or kJ mol^{-1}.
Electron affinity is represented by negative sign (-)

$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} + 340$ kJ/mol.

Therefore, electron affinity of chlorine is -340 kJ/mol.

<u>Successive electron affinities</u>.
When the first electron is added to an atom it forms a monovalent anion, however if the 2nd electron is added to the same anion, it experiences a repulsion and energy is absorbed during this process therefore the 2nd and further successive electron affinities are positive. Electron affinities depend upon atomic size and nuclear charge.

(a) the smaller the atomic size the greater the electron affinity because the attraction between the nucleus and the valence electrons is greater.
(b) The greater the nuclear charge, greater is the electron affinity because of more attraction on the valence electrons from the nucleus.



How does electron affinity vary across the period and down the group?

In a period As the atomic size decreases, the nuclear charge increases so the electron affinity increases.
Electron affinity for group 17 elements is the highest and the least is for group 1.
Down a group moving from top to the bottom of a group, the atomic size increases more than the nuclear charge therefore causing a decrease in the electron affinity.
Note: Fluorine has lower electron affinity than chlorine.
Inert gases have zero electron affinity due to their stable electronic configuration.
Group 2 and group 15 have fully filled and half filled sub-orbitals. The electron affinity for these groups does not show a negative value.

Explain Electronegativity:
Definition : The tendency of an atom in a molecule to attract the shared pair of electrons towards itself is called is electronegativity.
Electronegativity only indicates the tendency of different elements to attract the bond forming electron pair.
Electronegativity has the highest value for fluorine which is 4 and the lowest value for Caesium which is 0.7
Electronegativity values depend upon (a) Atomic size (b) Nuclear charge

(a) **Atomic size** : The greater the size of the atom, the lesser the electronegativity since the electrons are far away from the nucleus.
(b) **Nuclear charge** : More nuclear charge more attraction for electrons, greater the electronegativity.

How does electronegativity change across the period and down the grou[?

Across the period atomic size decreases, nuclear charge increases, electronegativity increases.
Down a group atomic size increases along with the nuclear charge but the effect of an increase in atomic size overcomes the effect of an increase in nuclear charge hence electronegativity increases.

Note: Generally metals show low electronegativity compared to non-metals. In other words, metals are electropositive and non-metals are electronegative.
Noble gases have a complete octet so they do not have a tendency to attract electrons.

Note:
(i) Elements with n/p (neutron/proton) ratio around 1 are stable. E.g. light metals like sodium postassium and calcium.
(ii) Elements with n/p ration 1.5 and above are radioactive, i.e. they emit radiations. They are unstable elements, e.g heavy metals like uranium.

COMPARISON OF ALKALI METALS AND HALOGENS

	Alkali Metals (Group 1)	Halogens (Group 17)
Elements	Li, Na, K, Rb, Cs, Fr	F, Cl, Br, I, At(coloured non metals)
Occurrence	Commonly found in a combined state due to their reactive nature	Commonly found in a combined state as salts.
Physical State	Shining white solid metals. Soft and can be cut with a knife. Lithium is the hardest.	Non-Metals : Diatomic in the gaseous state. Flourine (Gas) Chlorine (Gas), Bromine (liquid), Iodine (Solid).
Valence electrons	Possess one valence electron and therefore show similar properties.	Possess seven valence electrons each and therefore show similar properties.
Conduction	Good conductors of electricity.	'Non-conductors of electricity.
Nature	Highly reactive, electropositive metals. Metallic character, increases from lithium (Li) to francium (Fr).	Highly reactive, electronegative non-metals. Non-metallic character, decreases from fluorine (F) to iodine (I).
Melting point and boiling point	Decreases down the group.	Increases down the group.
Atomic size	They have the largest atomic size in their period (except inert gases).	They have the smallest atomic size in their period. The atomic size increases down the group.
Ionisation energy	They have lowest I.E. in their period. It decreases down the group.	They have high I.E. (lower than noble gases) in their period.
Electron affinity	They have low E.A. values which further decrease down the group.	They have high E.A. values. They too decrease down the group.
Electronegativity	They have the lowest E.N. in their period. E.N. decreases down the group.	They have high E.N. highest in their period.
Reactivity	They are reactive metals. Reactivity further increases down the group.	They are reactive non-metals. The reactivity of these non-metals decreases down the group.
Reaction with water and acids	They react vigorously with water and acids liberating hydrogen. Reactivity further increases down the group.	Generally they do not react with dil. acids and water.
Reducing/ oxidising agents	Strong reducing agents as they lose electrons to complete their octet.	Strong oxidising agents as they accept electrons to complete their octet.
Compound formation	Form electrovalent compounds with non-metals. Example : NaCl, KBr.	Form electrovalent compounds with metals. e.g., KCI, $CaC1_2$. Form covalent compounds with hydrogen and other non-metals, e.g., HBr, HCI, HI, CCl_4.

VARIATION IN PERIODIC PROPERTIES

Properties	Across the period (Left to right)	Down the group (Top to bottom)
Atomic size (radius)	Decreases	Increases
No. of valence electrons	Increases	Remains same
Metallic character	Decreases	Increases
Non-metallic character	Increases	Decreases
Electron affinity	Increases	Decreases
Electronegativity	Increases	Decreases
Ionisation energy	Increases	Decreases
Basic nature of oxides	Decreases	Increases
Melting point	Increases from group I to group IV and then decreases	Group I and II decreases Group III and IV decreases Group V to VII increases
Boiling point	Increases from group I to group IV and then	Group I and II decreases Group III and IV decreases
Oxidising nature	decreases	Group V to VII increases
Reducing nature	increases	Decreases

****