Periodic table

For 1st basic students science (faculty of education)

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History of periodic table

During the **nineteenth century**, Chemists have always looked for ways of arranging the elements to reflect the similarities between their chemical and physical properties, to known the elements application and its uses

1- Johann Dobereiner

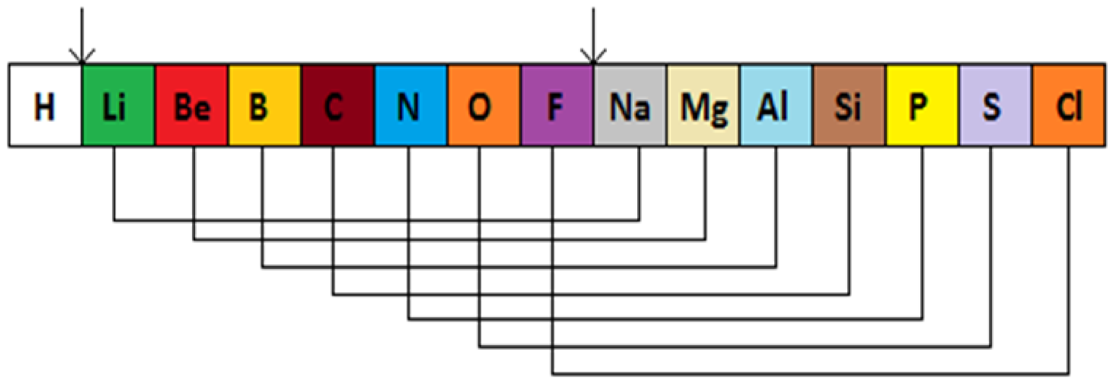
began to formulate one of the earliest attempts to classify the elements. In 1829, he found that he could form some of the elements into groups of three, with the members of each group having related properties. He termed these groups ***triads***. He found the middle element had properties that were an average of the other two members when ordered by the atomic weight (the Law of Triads)

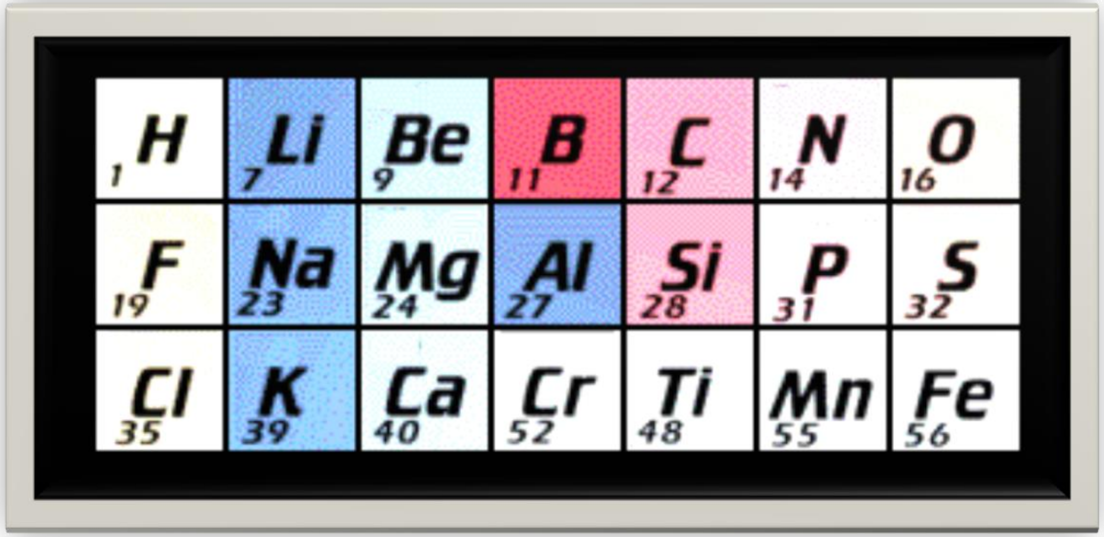
Alkali formers	
Li	7
Na	23
K	39

Salt formers	
Cl	35.5
Br	80
I	127

History of periodic table

2- John Newlands:. The atomic masses of many of the elements had been determined. He suggested that the chemistry of the elements might be related to their masses. **The known elements (>60) were arranged in order of atomic weights. He noticed that the chemical and physical properties repeated every 8th element when placed in order of increasing atomic mass. Observed similarities between the first and 8th elements, the second and 9th elements etc**





History of periodic table

2-Disadvantages of John Newlands

The problem was found heavier than Calcium the pattern starts to breakdown. Although Newland had the right idea.

1-That is due to some of the elements hadn’t been discovered yet and he did not leave any gaps for undiscovered elements in the table.

2- Sometimes had to cram two elements into one box in order to keep the pattern due to the similarity in the atomic weight.

But these elements have a different chemical and physical properties.

History of periodic table

3- Dmitri Mendeleev

Produced a table based on atomic weights but arranged 'periodically' with elements with similar properties under each other

3-Disadvantages of Dmitri Mendeleev

1- The ascending order of the atomic weights of some elements was destroyed.

I يود	Te تيلور	العنصر
126.8	128.3	الوزن الذري

History of periodic table

3-Disadvantages of Dmitri Mendeleev

Where the Tylor has the same properties of oxygen group and Iodide has a properties similar to chloride group so iodide and tylor must be place changed Regardless of the atomic weight value

3-Advantages of Dmitri Mendeleev

Mendeleev arranged 63 elements in a table leaving spaces for elements that had not been discovered and predicted these elements such as scandium, gallium, and germanium later, with determination their approximate atomic balance

History of periodic table

4- Henry Moseley

1-Henry Moseley discovered that the positive charge in the nucleus of an atom of any element is of a definite amount.

2-Determined the atomic number of each of the elements

3-He arranged the elements in order of increasing atomic number (nuclear charge).

The differences in Mendeleev's table and the modern periodic table

History of periodic table

The differences in Mendeleev’s table and the modern periodic table

1-Mendeleev’s table was arranged in order of increasing atomic mass. Modern table is arranged in order of increasing atomic number.

2-In Mendeleev’s table the noble gases are not included but included in the modern table

3-There are gaps in Medeleev’s table but there are’t in the modern periodic table as they have been discovered..

Periodic Table of the Elements

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
IA	IIA											IIIA	IVA	VA	VIA	VIIA	VIIIA
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.41	31 **Ga** 69.72	32 **Ge** 72.64	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (97.9)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
55 **Cs** 132.9	56 **Ba** 137.3	57 **La*** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.8	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 **Ac~** (227)	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Uuu** (272)	112 **Uub** (277)	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**		

***Lanthanides**	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
~Actinides	90 **Th** 232.0	91 **Pa** (231)	92 **U** (238)	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

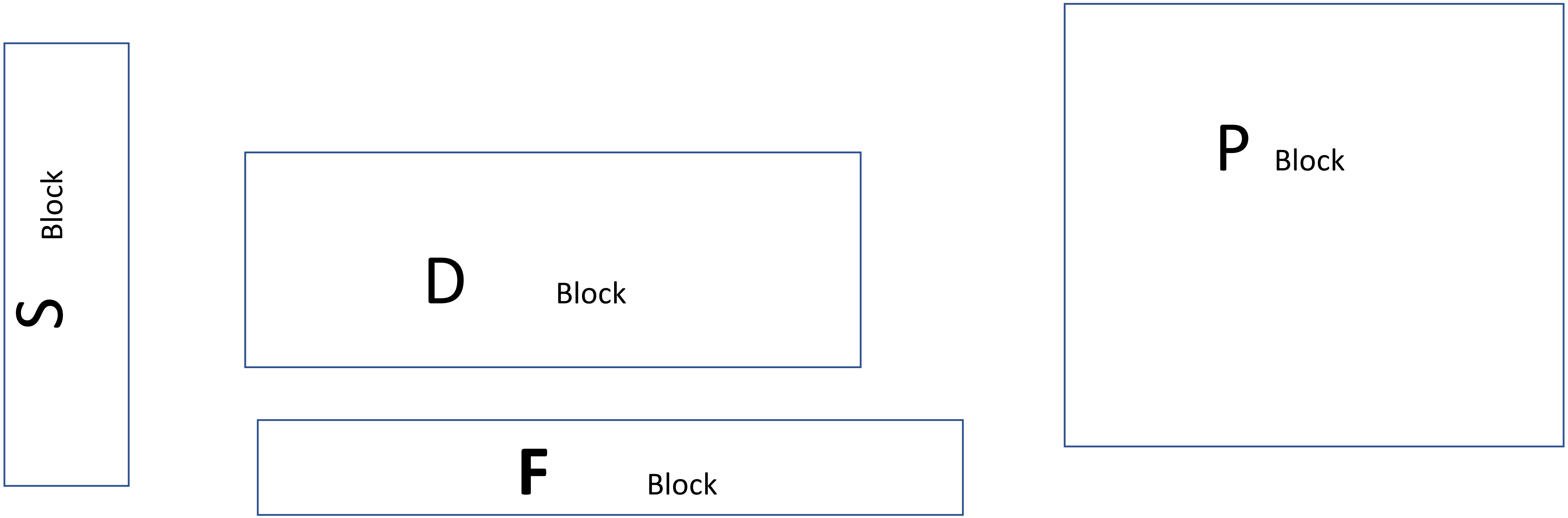
Description of periodic table

Modern periodic table consist of:.

1- 18 group vertical

2- 7 rows horizontal

And can be divided into 4 blocks



Gradation of proprieties in periodic table

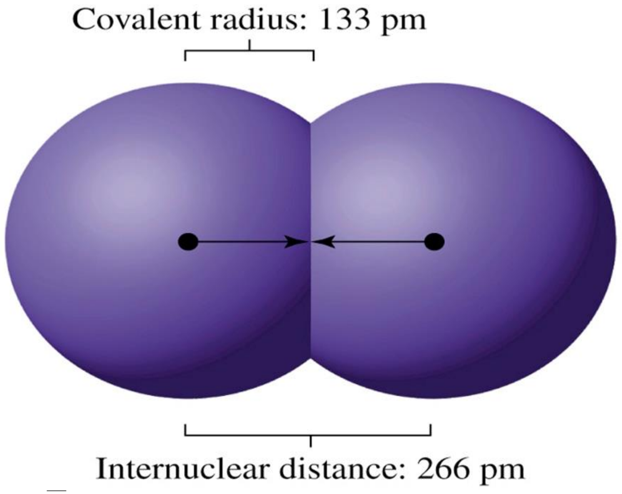
1- Atomic radius.

The distance between the center of two identical atoms in one molecule.

A- In row:

By increasing the atomic number, the atomic size decrease

that is Due to:. increasing the positive charge in the nucleus and the electron added in the same level So there are't any further shield from the nucleu attraction



Gradation of proprieties in periodic table

1- Atomic radius.

The distance between the center of two identical atoms one molecule.

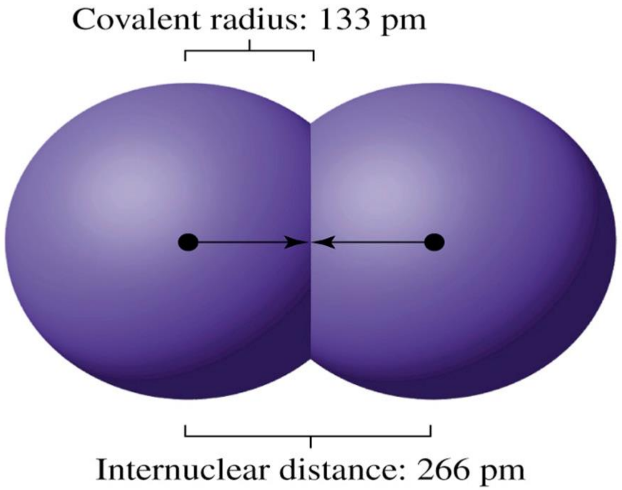
B- In column:

By increasing the atomic number the atomic size increase

that is Due to:.

1-the electron added in the new level far away from the nucleus charge

2-Full Energy levels are mask (shield) the effect of the nucleus force form the outer electrons level



Gradation of proprieties in periodic table

2- Ionic size.

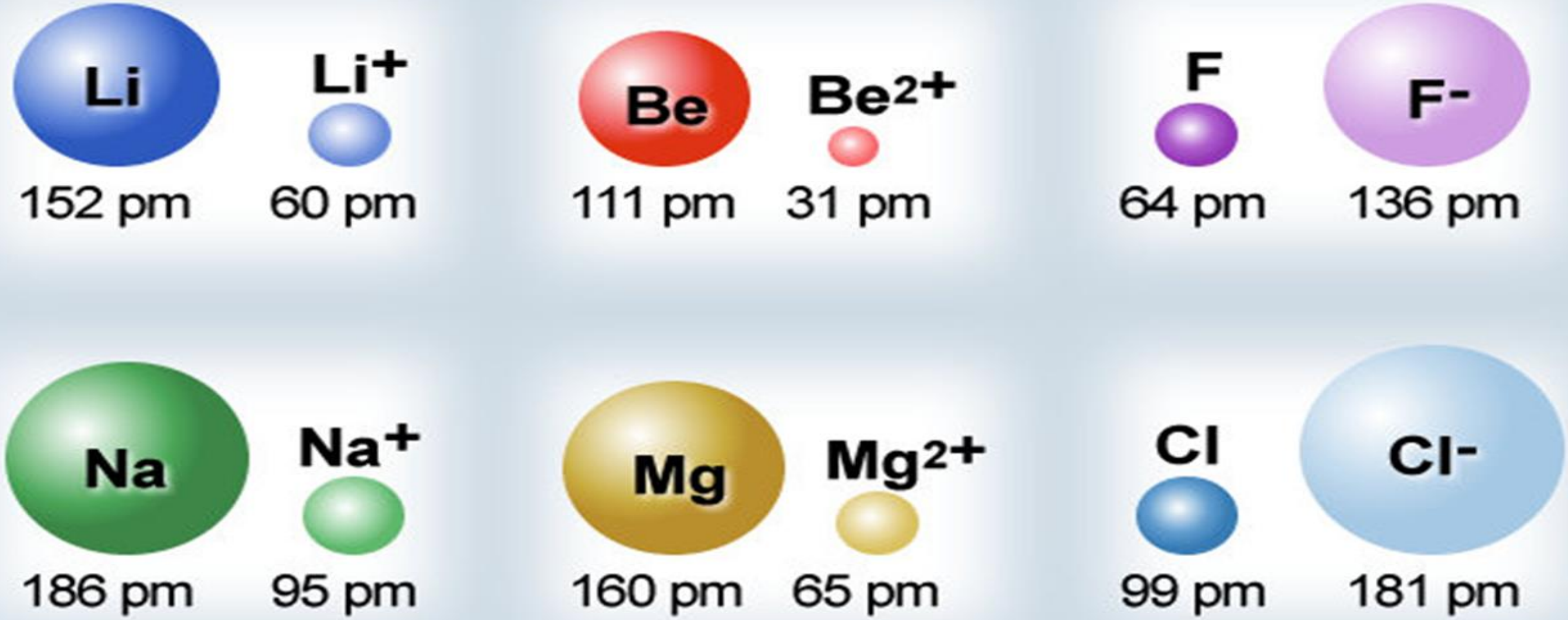
A- Positive ion:

In positive ion the outer electrons removed so the nucleus now has more positive charge than the total negative charge (electrons). The larger effective nuclear charge will now pull electrons closer to the nucleus

B- Negative ion:

The Nucleus charge remain the same with increasing the number of electrons leading to1- increasing the repulsion force between the electrons with each other that is caused an increase in the energy level radius

2-The effect of the nucleus charge on the electrons attraction involved decreases


Li
Li+
152 pm
60 pm
Be
Be2+
111 pm
31 pm
F
F-
64 pm
136 pm
Na
Na+
186 pm
95 pm
Mg
Mg2+
160 pm
65 pm
Cl
Cl-
99 pm
181 pm

Radii

Gradation of proprieties in periodic table

3-Ionization Energy (IE)

The amount of energy needed to remove an electron from an atom or ion. Every electron in any atom or ion has a specific ionization energy.

<u>Ionization Energy Trends</u>

<u>1) Ionization energy DECREASES going DOWN a GROUP</u>

As we move from up to down in a group, the charge of the nucleus decrease and the radius increase. Leading to decrees the attraction force between the nucleus and the outer electron, so it does not requires a high energy to separate it from the atom

Gradation of proprieties in periodic table

3-Ionization Energy (IE)

2) Ionization energy INCREASES as you go ACROSS a PERIOD

As we move from left to right in a period, the charge of the nucleus increases and the radius decreases. Leading to an increase in the attraction force between the nucleus and outer electron so it requires a large energy to separate it from the atom until it reaches the maximum value in the noble gases

3) The noble gases have the highest Ionization energies

4) The alkali metals have the lowest Ionization energies

Gradation of proprieties in periodic table

3-Ionization Energy (IE)

<u>First Ionization Energy</u>

The amount of energy needed to remove the 1st electron from the outer shell of a neutral (uncharged) atom.

- **Some atoms have 2 or more electrons in the valence Shell. For example, alkaline earth elements such as Mg or Sr**
- **Removing the 1st electron (first ionization energy) will always take less energy than the energy necessary to Remove the 2nd electron (second ionization energy).**
- **This trend continues. The energy necessary to remove a 3rd electron is greater than the energy necessary to Remove the 2nd and so on <u>that is due to Once an e^- is removed the nucleus charge per electron is now stronger. So the I.E. to remove the next electron increases</u>**

Gradation of proprieties in periodic table

<u>Second ionization energy</u> is the energy required to remove the second electron from ion +1

<u>Third ionization energy</u>,, is the energy required to remove the third electron from ion +2

<u>4-Electronegativity</u>

The ability of atom to attract the electrons towards it <u>in a chemical bond</u>

1) **Electronegativity decreases down a Group**
2) **Electronegativity increases across Periods**

As I.E

Generally, metals are electron givers and have low electronegativities. Nonmetals are electron takers and have high electronegativities

Gradation of proprieties in periodic table

5-Electron affinity

Energy change that occurs when an atom gains an electron

<u>In period</u> :As we move from left to right in a period, the electronic tendency increase as the radius decrease and the force of attraction of the nucleus increases

<u>In group</u>: the electron tendency decreases vertically in the group by increasing the atomic number because electron shielding blocks some of the attraction force from the nucleus, and the radius of the atom increases

<u>Irregular cases:</u> The irregularity in beryllium because its sub-levels are empty and the atom is stable. In nitrogen, the sub-levels are half-full as in the case of the neon atom. All sub-levels are full.

Gradation of proprieties in periodic table

6-Mineral and non-metallic properties:

In period : The period begins with a strong metal element that gradually reduces the metal content by increasing the atomic number to reach the semi-metals then The non-metals are shown by increasing the atomic number

In group: The metal is increased by increasing the atomic number, because the higher the atomic number, the greater the number of energy levels

Semi-metal : Elements whose characteristics range from the properties of ordinary metals and non-metals are called semiconductors. Boron, silicon, germanium, arsenic, antimony, and tellurium are examples of these elements. Semiconductors are characterized by the following :

- **The possibility of gaining or losing electrons during the reaction.**
- **Usually connected to electricity and heat, but not to the quality of metal elements and its faintly color.**

Gradation of proprieties in periodic table

Non-metals	Metals
Are elements that fill their outer shell with more than four electrons	Are elements that fill their outer shell with less than four electrons
The atom tends to gain electrons in its outer shell and become negative ions to reach the nearest inert gas structure	The atom tends to lose electrons in its outer shell and become positive ions to reach the nearest inert gas structure
Non-metals are bad conductors of heat and electricity.	Metals are good conductors of heat and electricity.
Characterized by a small radius	Characterized by a large radius
Non-metals have low melting and boiling points as compared to metals.	The melting and boiling points of metals are generally high

Gradation of proprieties in periodic table

Non-metals	Metals
Non-metals are not strong.	All metals are strong.
Most of Non-metals are soft	Generally, metals are hard.
Non-metals are non-malleable or brittle	Metals are malleable.
It has no metallic luster.	It has a metallic luster.
High electronegativity, High ionization energy, high electron affinity, and lower electro-positivity	lower electronegativity, lower ionization energy, lower electron affinity, and high electro-positivity

General properties of Hydrogen

- <u>Hydrogen: </u>Is the first element in the periodic table and characterized by the most simple form of the atom, where the nucleus contains a single proton and one electron can spin around it

- <u>Hydrogen is characterized.</u>
  - There is no really appropriate place for this element in the periodic table but it placed in group A1. But it is considered a class by itself. It is associated with alkaline metals and halogens.
  - Hydrogen is therefore similar to alkaline metals which contain a single electron that can be lost and became ionized proton.

General properties of Hydrogen

- <u>There is a difference between hydrogen and the elements of group A1 in the following:.</u>
- 1 - Hydrogen has a small potential to lose an electron In contrast to the alkali elements can lose an electron easily and therefore the ability of hydrogen to form ionic compounds is difficult
  2 - Hydrogen has great potential to form covalent compounds by electron participation

<u>Its relationship with the halogen group:</u>

- The hydrogen atom contains only one electron and it needs one electron to reach the structure configuration of the nearest inert gas which is helium.

General properties of Hydrogen

- The same case for halogens need one electron to reach the structure of the nearest inert
  With the electrical analysis of the LiH (Hydride) solution we found that hydrogen is released to the elevator and this is similar to the electrolysis of the NaCl solution. Chlorine is released into the elevator as well.
- <u>There is a difference between hydrogen and the elements of group A7 in the following:.</u>
- The chance to form hydride is more difficult than the formation of halide that is due to the small electronegativity and electron affinity of hydrogen

Preparation of Hydrogen

- **<u>First: Displacement of hydrogen from its compounds</u>**
- $2Na+H_2O \longrightarrow NaOH+H_2$
- $Zn+2HCl \longrightarrow ZnCl+H_2$
- $Zn+2NaOH \longrightarrow Na_2ZnO_2+H_2$
- $NaH+H_2O \longrightarrow NaOH+H_2$
- $CaH_2+H_2O \longrightarrow Ca(OH)_2+H_2$
- **<u>Second: Electrolysis:</u>**
- Hydrogen is produced through electrolysis of water, which is the most common method of producing hydrogen in relatively low cost and does not produce pollutants to the environment during the process
    - $H_2O + \text{electricity} \rightarrow H_2 + 1/2\ O_2$

Reaction of Hydrogen

➢ Reduction Reactions of Hydrogen

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

$$H_2C{=}CH_2(g) + H_2(g) \rightarrow H_3C{-}CH_3(g)$$

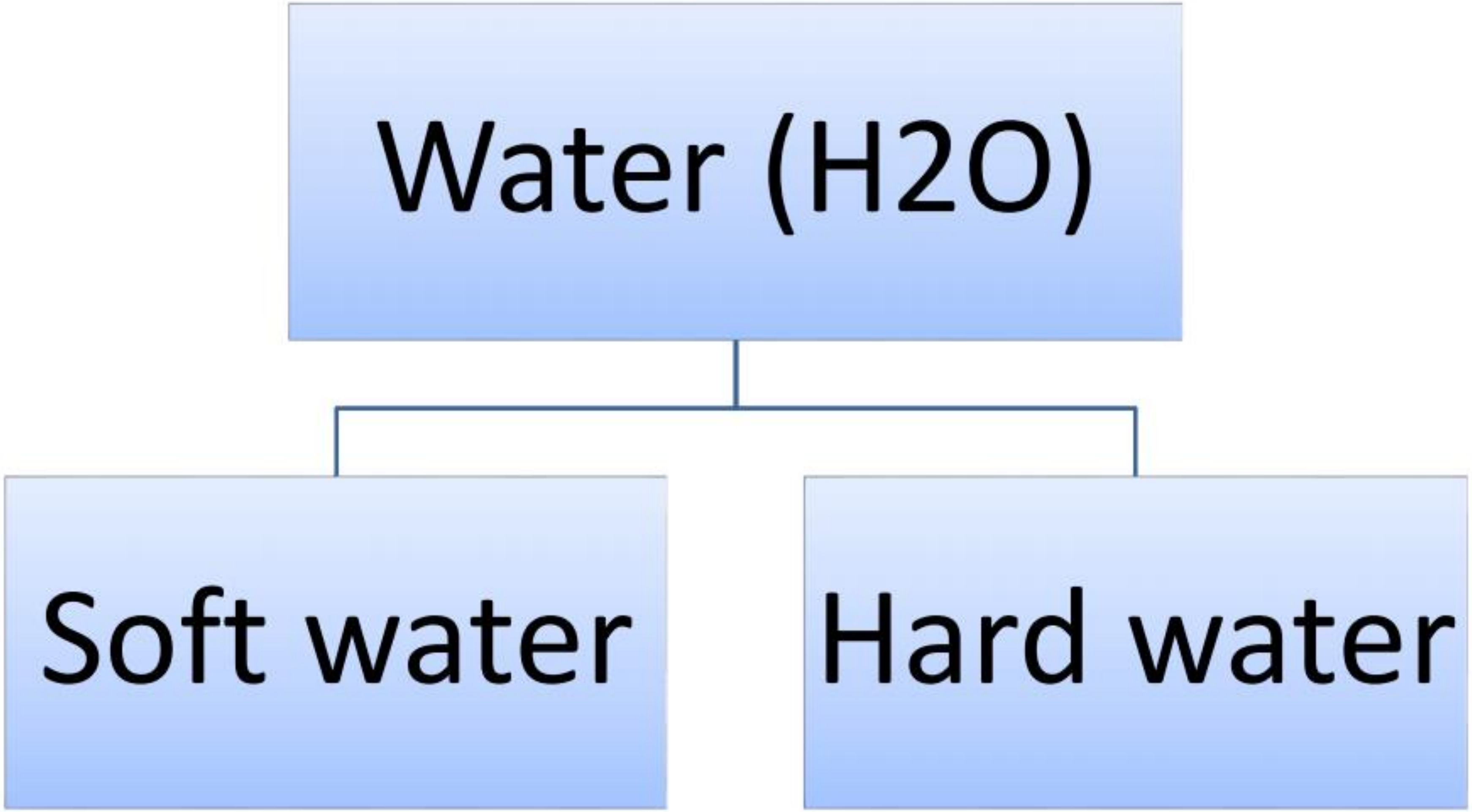
➢ another Reactions of Hydrogen

$$2H_2 + O_2 \rightarrow 2H_2O$$

$$N_2 + 3H_2 \rightarrow 2NH_3$$

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

Hardness of water



Hardness of water

- **Soft water**

Water consists of low concentration of calcium and magnesium salts.

- **Hard water**

Hard water is due to the presence of high concentration of calcium and magnesium salts that are dissolved in water

Temporary Hardness

Is due to the presence of Ca^{2+} , Mg^{2+} in the form of the bicarbonate ion HCO $^{3-}$, being present in the water.

This type of hardness can be treated by boiling the water to expel the CO_2, as indicated by the following equation:

.

- $Ca^{2+} + 2HCO_3^- \rightarrow CaCO_3 + H_2O + CO_2$
- $Mg^{2+} + 2HCO_3^- \rightarrow MgCO_3 + H_2O + CO_2$

Hardness of water

<u>Permanent hardness</u> is due to the presence of the ions Ca^{2+} , Mg^{+2} in the form of Cl^{-} and SO_4^{2-}. This type of hardness cannot be eliminated by boiling

As it can't be treated easily ,so it's treated by chemical treatment such as: CLARK'S METHOD ion exchange resin.

<u>CLARK'S METHOD</u>

In clark's method an amount of calcium hydroxide $Ca(OH)_2$ and sodium carbonate is added to hard water. Which convert soluble calcium to insoluble carbonates

$$CaSO_4 + Na_2CO_3 \longrightarrow Na_2SO_4 + CaCO_3$$

.

Hardness of water

<u>Ion Exchange resin</u>

In this method sodium zeolite is used as ion-exchanger. Na-zeolite is passed through the pipes containing hard water. Sodium zeolite is converted into calcium-zeolite or magnesium-zeolite. These are insoluble in water and are separated from water by filtration

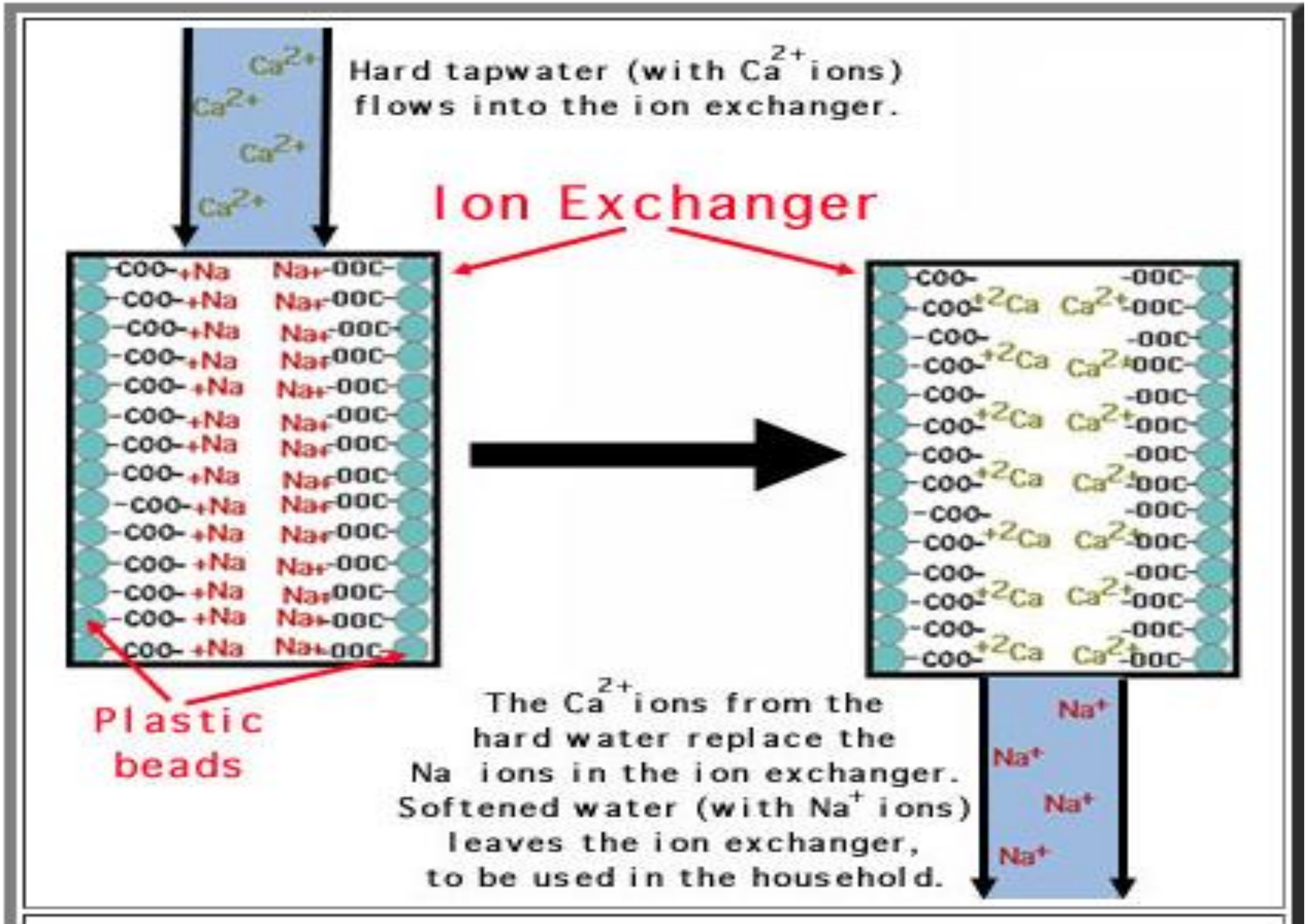
$$Ca^{+2} + Na_2\text{-zeolite} \longrightarrow Ca\text{-zeolite} + 2Na^{+1}$$

$$Mg^{+2} + Na_2\text{-zeolite} \longrightarrow Mg\text{-zeolite} + 2Na^{+}$$

A zeolite which has been used can be regenerated by allowing it to stand in contact with conc. NaCl solution. The calcium aluminosilicate is reconverted to sodium aluminosilicate:

$$Ca(AlSi_2O_6)_2 + 2Na^+ \rightarrow 2NaAlSi_2O_6 + Ca^{2+}.$$

Hardness of water



Hardness of water

Disadvantages of Hard water

Domestic: Hard water affect cleaning ability of soap. When hard water is used for washing, large amount of soap is consumed.

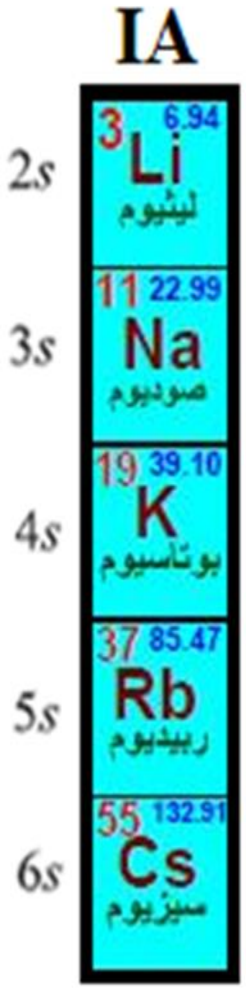
(ii) It is unsuitable for cooking certain vegetables and meat. They take very long time to cook in hard water

Industrial: Hard water can cause “corrosion" inside the pipes that transport water .

Health: Hard water when used for drinking for long period can lead to stomach and kidney diseases.

1^{st} group
Alkali Metal Properties

- The alkali family is found in the first column of the periodic table
- Atoms of the alkali metals have a single electron in valence shell (1 valence electron) and electron configuration of $1s^1$.
- Tend to lose an electron and turn to positively charged ions, carrying one positive charge M^{+1}
- It is characterized by the largest size of all the atoms in the periodic table and the atomic volume increases from top to bottom and according to this characteristic, these elements are:
- They are the most reactive metals, and characterized by having high electropositive properties that is due to the electron in its outer shell in these metals is easily lost



1^{st} group
Alkali Metal Properties

- So it is kept under the surface of kerosene or paraffin to prevent its interaction with the wet air because it is chemically active elements with hydrogen evaporation that ignites with a bang. Soft metals (easily cut with a knife).
- Alkali metals are never found as free elements in nature. They are always bonded with another element
- All the ions of these elements are smaller than the size of the corresponding atoms, because the positive charge inside the nucleus is greater than the number of electrons. This results in an increase in the nucleus's attraction force to the electrons and the electrons become more correlated with the nucleus.
- These elements have low electronegativity properties and therefore tend to form ionic compounds

1st group
Alkali Metal Properties

- **These elements are easily oxidized and converted into positive ion thus these elements considered powerful reducing agents. {Oxidation process: the process of electron loss accompanied by an increase in positive charge and a decrease in negative charge}**
- **The melting point of its elements is low and decreases from top to bottom because the metal bond in these metals is composed of one free electron per atom. The metal bond is weak and became more weaker by increasing the atomic number.**
- **They have low melting and boiling points and are soft and easy to see. Lithium can be easily sliced with a knife while cesium is as soft as cheese.**
- **Lowest ionization energies and electronegativities in periodic table, easily ionized to form ions with +1 charge/+1 oxidation state**
- **They are good conductors of electricity.**
- **Their ability to conduct electricity is due to the availability of their outer electrons to move.**

1st group
Alkali Metal Properties

- The solutions of these elements are not colored because the resulting ions do not contain free electrons
- Do not prepare in pure form by reducing their ions but they are brought by electrolysis of their salts.

1^{st} group
Alkali Metal Properties

- **<u>Chemical reactions:.</u>**
- **<u>1-Oxides of the alkali metals</u>**

○**Alkali metal *on combustion in Oxygen***

$M+O_2 \rightarrow$depends on the alkali metal

Li_2O(oxide)

Na_2O_2(peroxide)

KO_2,RbO_2,CsO_2(superoxide)

The increasing stability of the peroxide or superoxide, as the size of the metal ion increases, is due to the stabilization of large anions by larger cations through lattice energy effects

1^{st} group
Alkali Metal Properties

- **<u>Chemical reactions:.</u>**
- **<u>1-Oxides of the alkali metals</u>**

- **All oxides compounds react violently with H_2O (hydrolysis) forming hydroxide**

$$Li_2O(s) + H_2O(l) \longrightarrow 2\,LiOH(s)$$

$$Na_2O_2(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(s) + H_2O_2(l)$$

$$2\,KO_2(s) + 2\,H_2O(l) \longrightarrow 2\,KOH(s) + H_2O_2(l) + O_2(g)$$

The alkali metal hydroxides are the strongest of all bases.

- **<u>2-Reaction with water:</u>**
- **All alkali metals react with water to produce hydrogen and the metal hydroxide.**

$$M(s) + H_2O(l) \longrightarrow M^+\,(aq) + OH^-\,(aq) + \tfrac{1}{2}\,H_2(g)$$

1^{st} group
Alkali Metal Properties

- **Chemical reactions:.**
- **3-Reaction with halogens (group 17: F_2, Cl_2, Br_2, I_2)**

All alkali metals react with halogens to form ionic halides.

$$M_{(s)} + 1/2\ X_2 \longrightarrow MX$$

- **4-Reaction with hydrogen:.**
- $M_{(s)} + 1/2\ H_2 \longrightarrow MX_{(s)}$

1^{st} group
Alkali Metal Properties

- **<u>Points of Difference between Lithium and other Alkali Metals</u>**
- **Exceptionally small size of its atom and ion**
- **Lithium is much harder than the other elements in group.**
- **Its m.p. and b.p. are higher than the other alkali metals.**
- **it is activity is smaller than other elements in this group**
- **On combustion in air it forms mainly monoxide ,Li_2O and the nitride , Li_3N unlike other alkali metals.**

2^{nd} group
Alkaline earth metal Properties

Elements whose atoms have their s-subshell filled with two valence electrons are called alkaline earth metals. Their general electronic configuration is [Noble gas] ns^2.
They occupy the second column of the periodic table and so-called as group second groups.

Group 2A (or IIA) of the periodic table are the alkaline earth metals: beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).

- **They are harder and less reactive than the alkali metals of Group 1A. The name comes from the fact that the oxides of these metals produced basic solutions when dissolved in water.**
- **Like the Group 1A elements, the alkaline earth metals are too reactive to be found in nature in their elemental form.**

. The alkaline earth metals have two valence electrons in their highest-energy orbital's (ns2). They are smaller than the alkali metals of the same period, and therefore have higher ionization energies. In most cases, the alkaline earth metals are ionized to form a 2+ charge.

2^{nd} group
Alkaline earth metal Properties

Atomic and Ionic Radii

Ionic and Atomic radius increases down the column of the periodic table, both radii will be smaller than the alkali metal and larger than other atoms of the same period.

Why Alkaline Earth Metals are Denser than Alkali Metals?

Radii being smaller, the volume of the atoms are also smaller. In addition, due to the presence of two valence electrons, atoms have stronger metallic bonding. Hence, alkaline earth metals have more density and harder than alkali metals.

Ionization Energy

Alkaline earth elements can donate both valence electrons to get a noble gas configuration of octet configuration. Thus, they have two ionization energies:

First Ionization energy

The first ionization energy of alkaline earth metals is the energy needed for the removal of the first electron from the neutral atom. It is larger than that of the alkali metal atom for

Due to smaller radii and the electrons being held tightly by the higher nuclear charge

2^{nd} group
Alkaline earth metal Properties

Second Ionization Energy

The second ionization energy of alkaline earth metals needed for the second electron from the cation will be more than the first ionization energy of the atom.

So group two alkaline earth elements are all divalent electropositive metals and exhibit a fixed oxidation state of 2. Ionization energy needed for the removal of the valence electron will be highest for the small beryllium atom.
With increasing atomic size, the valence electron gets shielded by the inner electrons and becomes easily removable with less energy requirement. Hence the ionization energy decreases with an increasing atomic number or atomic size.

Reactivity of Alkaline Earth Metals

Reducing ability is inversely related to ionization energy. As ionization energy decreases down the column, reducing property is expected to increase from Beryllium to Barium.
Reduction potential also decreases from beryllium to barium indicating the increasing reducing capacities. But, the alkaline earth metals are weaker reducing agents than alkali metals, due to higher ionization energy.

2^{nd} group
Alkaline earth metal Properties

Melting and Boiling Points

Because of smaller size and strong metallic bonding in close-packed structure, the melting and boiling points of the alkaline earth metals are higher than alkali metals. the melting and boiling points decrease regularly from beryllium to barium.

Anomalous behaviour of Beryllium

Beryllium has more covalent nature due to its smallest size, Highest ionization energy, low electropositive nature. Because of these, Beryllium differs from other alkaline earth metal properties.

It is the hardest metal among alkaline earth metals

Does not react with water even at red hot conditions.

Melting and boiling point of beryllium is maximum.

It does not react directly with hydrogen to form hydride.

Uses of the Alkali Earth Metals

Magnesium usually burns with a bright whitish flame and this has allowed it to be used in fireworks. Since calcium forms such hard minerals, it is useful in building materials, such as plaster, mortar, and cement. Mortar is made from calcium oxide, CaO, also known as lime, or quicklime. When calcium oxide is treated with water it forms calcium hydroxide, $Ca(OH)_2$, or slaked lime, which absorbs carbon dioxide from the air and gradually forms calcium carbonate, $CaCO_3$.

3rd group
Boron group

5 B
13 Al
31 Ga
49 In
81 Ti

Group 3A (or IIIA) of the periodic table includes the metalloid boron (B), as well as the metals aluminum (Al), gallium (Ga), indium (In), and thallium (Tl). Boron forms mostly covalent bonds, while the other elements in Group 3A form mostly ionic bonds.
The Group 3A metals have three valence electrons in their highest-energy orbitals (ns^2p^1). They have higher ionization energies than the Group 1A and 2A elements, and are ionized to form a 3+ charges.

This group characterized by irregularity in the physical properties. That is due to the appearance of d and f orbital.

General properties of the group
Ionization energies

The erratic way in which ionization energies vary among the elements of the group is due to the presence of the filled inner *d* orbital in gallium, indium, and thallium, and the *f* orbital in thallium, which do not shield the outermost electrons from the pull of the nuclear charge as efficiently as do the inner *s electrons.*

3^{rd} group
Boron group

As we know the ionization energies of these Group-1 and Group-2 elements decrease smoothly down the group. The ionization energies of gallium, indium, and thallium are thus higher than expected from their Group 2. that is due to the outer electrons, being poorly shielded by the inner *d* and *f* electrons, causind more strongly bound to the nucleus. This shielding effect also makes the atoms of gallium, indium, and thallium smaller than the atoms of their Group 1 and 2 neighbours by causing the outer electrons to be pulled closer toward the nucleus.
The M^{3+} state for gallium, indium, and thallium is energetically less favourable than Al^{3+} because the high ionization energies of these three elements. So the common oxidation state for these elements are +1

Atomic radius
By increasing the atomic number the atomic radius increase

Elecropositivity
Increase from B to Al and decrease again after Al

3^{rd} group
Boron group

Reaction of 3^{rd} group

1 -with oxygen

- All 3^{rd} react with oxygen to give the formula M_2O_3

$$M + O_2 \longrightarrow M_2O_3$$

Alumina oxide consider an inert oxide which can be used as cook coating.

2-with halogen

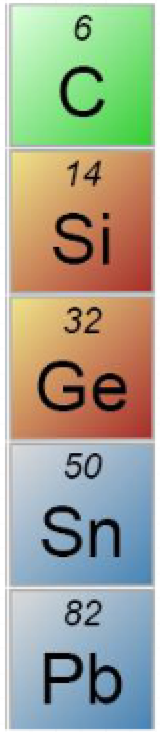
$$M + Cl_2 \longrightarrow MCl_3$$

3-with nitrogen

$$2M + N_2 \longrightarrow 2MN$$

4^{th} group
Carbon group

Carbon group elements



→ non-metal

→ metalloids

→ metals

Electronic configuration

$_{6}C$	$1s^2 2s^2 2p^2$	$[He]2s^2 2p^2$
$_{14}Si$	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
$_{32}Ge$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$	$[Ar]\ 4s^2 4p^2$
$_{50}Sn$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^2$	$[Kr]\ 5s^2 5p^2$
$_{82}Pb$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6s^2 6p^2$	$[Xe]6s^2 6p^2$

4^{th} group
Carbon group

The Group IV elements

→ outermost shell electronic configuration of ns^2np^2

- Elements of this group also exhibit multiple oxidation states.

- ***Lower*** oxidation states become more prominent down the group.....**Why?**

- ***Carbon*** forms predominantly covalent bonds, but the larger members of the group form ionic bonds with increasing ionic character.

4^{th} group
Carbon group

Physical and chemical properties of the group 4A elements

Like all other elements, these elements also follow a trend and a set patter of electronic configuration especially in their valance shells. They have 4 electrons in their valence shells.

Some of the chemical properties of group 4A elements are:

1-Atom of these elements has s2p2 configuration in the ground state.

2-They usually form covalent bonds to complete their octet.

3-The tendency to lose electrons increases on moving down the group as the size increases.

4-Metallic character increases on moving down the group.

5-Carbon forms negative ion carbide i.e. C4-. And silicon can form -4 forming silicide

6- Carbon shows +4 (as in CCl4)

7-Germanium and silicon are metalloids i.e. they have characters of both metals and non-metals.

8-Germanium and silicon can form +4 ions.

9-Lead and tin are metals.

10-Lead and tin can form +2 and unstable +4) due to the inert pair electrons

11-The atomic radii of the elements of 4A tend to increase on moving down the group as the atomic number increases.

4th group
Carbon group

Physical properties:

- Down the group, there is a regular decrease in m.p. and b.p. as the size of atoms increases and interatomic forces of attraction decreases.
- Hence silicon is hard while lead is a soft metal.

Atomic Radii (Covalent Radii)

- Atomic radii of these elements regularly increase as we move down the group primarily due to the addition of a new energy shell at each succeeding element.

Ionization energy

- In general, the ionization enthalpy decreases down the group. A small decrease in IE, from Si to Ge to Sn and a slight increase in IE, from Sn to Pb is the consequence of the poor shielding effect of intervening d and f orbitals and increase in the size of the atom.

Applications of 4A group elements
Carbon is used in pencils in the form of graphite.
Diamond is another form of carbon is used in jewelry.

Silicon is used in toothpastes, construction fillers, and glass.
Silicon is used is semiconductors.

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