General Chemistry
Chem101

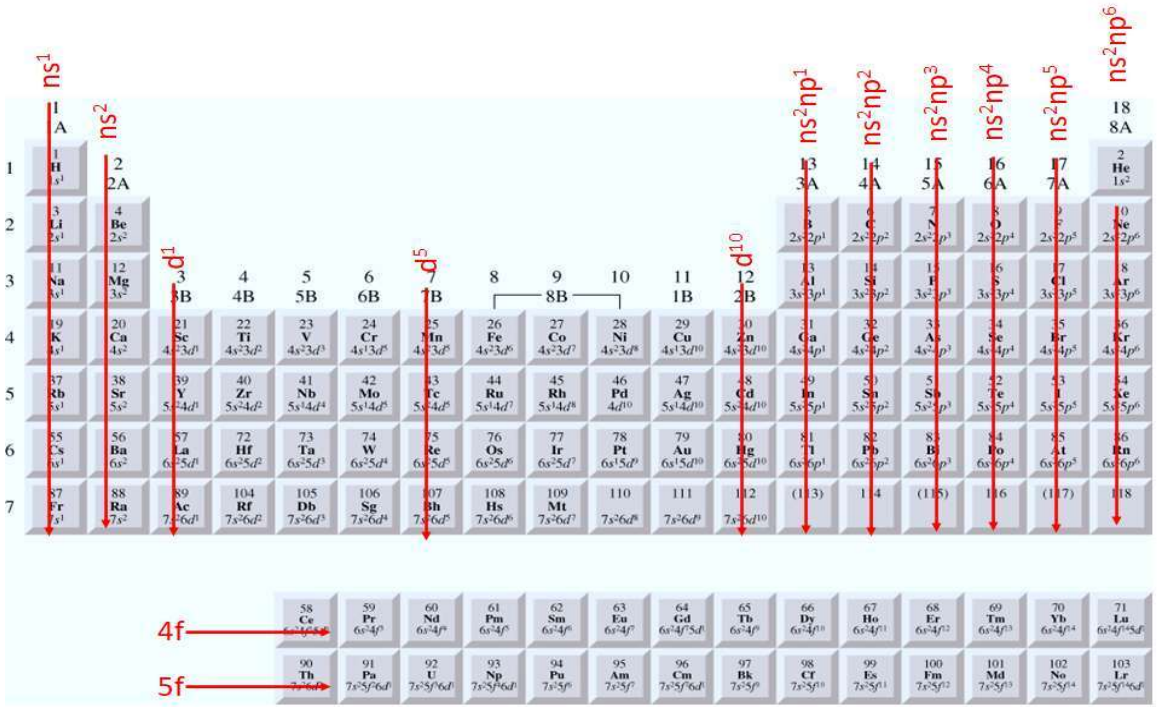
Chapter 3
Periodic Table

❖ Development of the Periodic Table

- Many attempts have been made to arrange the elements.
- When Russian Chemist Dmitri Mendeleev arranged the elements in order of increasing atomic mass, the properties were repeated. Because the properties repeated themselves regularly/periodically on his chart, the system became known as the periodic table.
- Mendeleev's classification was based on the atomic mass, however, this resulted in some inconsistency.
- Later on, after the discovery of the atomic number by Henry Moseley the **elements were arranged by their atomic numbers.**
- Periodic table is divided into four roughly rectangular areas called **blocks**. The rows of the table are called **periods**, and the columns are called **groups**.

Note:

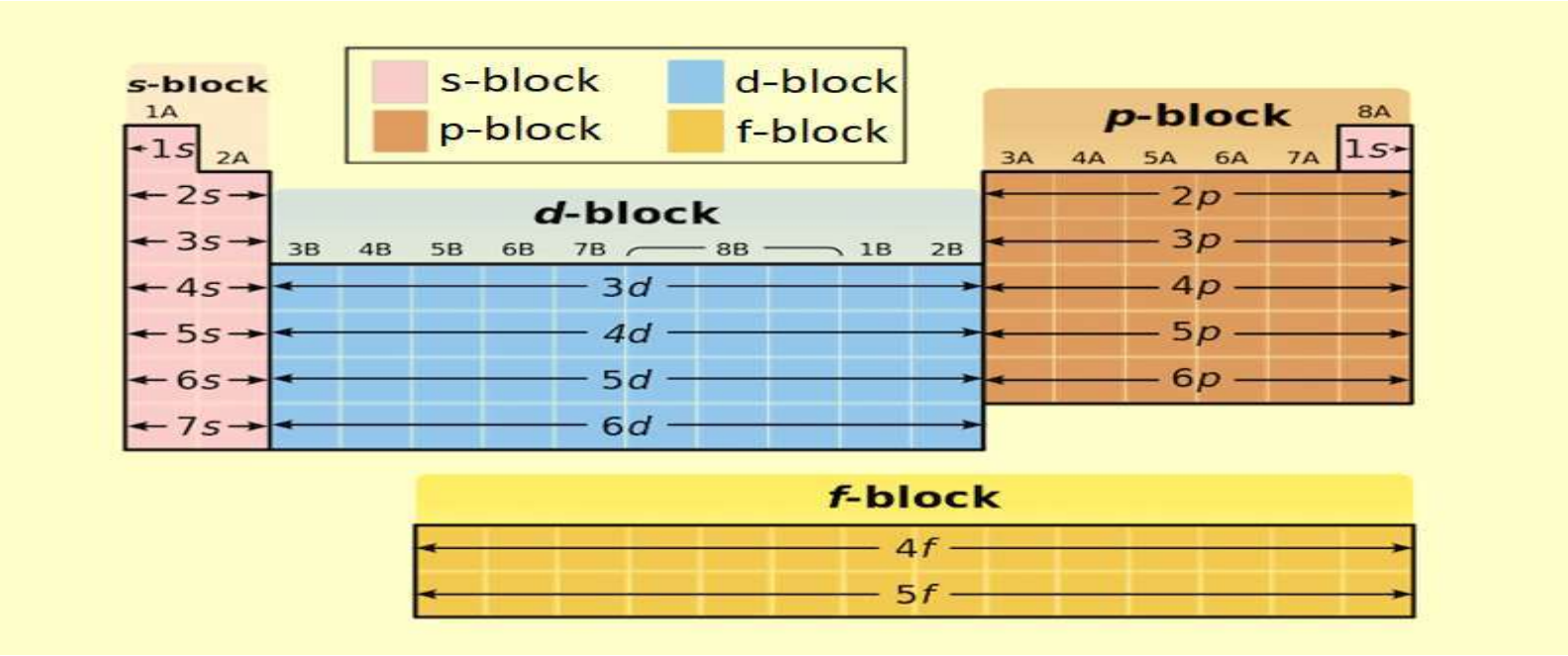
- The number of electrons increases as we go from left to right across the period.
- The value of n increases as we move down the group (from top to bottom).



1. Blocks of the Periodic Table| s-p-d-f Blocks

The periodic table is divided into **four** blocks depending on the type of shells being filled.

- **s-block** is characterized by **alkali metals** and **alkaline earth metals**.
- **p-block** is characterized by **metals, non-metals, and metalloids** of groups 3 to 8.
- **d-block** is characterized **by metals with variable (transition)** oxidation states.
- **f-block** is characterized by **Lanthanide and Actinide metals**.



How to find the block for an element in the periodic table ?

1. write the electronic configuration of the element
2. The orbital which receives the last electron represents the block of the element

Examples:

$_{3}Li: 1s^2\ 2\underline{s}^1 \Longrightarrow$ **s-block**

$_{6}C: 1s^2\ 2s^2\ 2\underline{p}^2 \Longrightarrow$ **p-block**

$_{21}Sc: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3\underline{d}^1 \Longrightarrow$ **d-block**

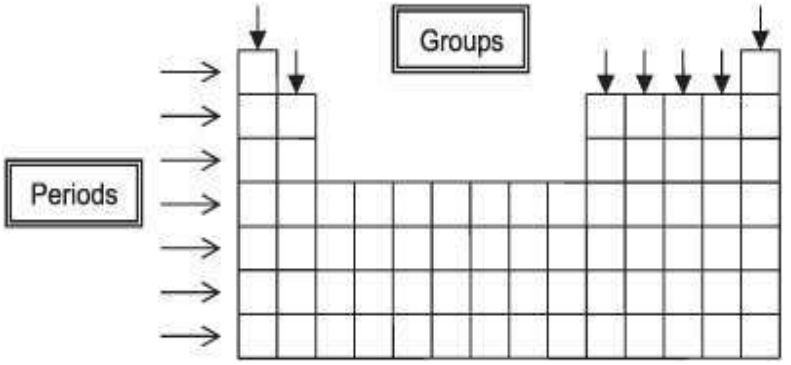
$_{58}Ce: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}5p^6\ 6s^2\ 4\underline{f}^2 \Longrightarrow$ **f-block**

2. Periods of the Periodic Table

- A **period** in the periodic table is a row of chemical elements
- There are currently **seven complete periods** in the periodic table

How to find the period for an element in the periodic table ?

1. write the electronic configuration of the element
2. the number of the period is equal to the highest principal quantum number (n) of the elements



➢ Period 1 (K)

- Highest principal quantum number **n=1**
- contains only two elements hydrogen and helium

Examples: $_{1}H:\underline{1}s^{1}$ $_{2}He:\underline{1}s^{2}$

➢ Period 2 (L)

- Highest principal quantum number **n=2**
- contains the elements lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine, and neon.

Examples: $_{3}Li : 1s^{2}\ \underline{2}s^{1}$ $_{6}C : 1s^{2}\ 2s^{2}\ \underline{2}p^{2}$

➢ Period 3 (M)

- Highest principal quantum number **n=3**
- contains 8 elements beginning with sodium and ending with argon.

Example: $_{11}Na : 1s^{2}\ 2s^{2}\ 2p^{6}\ \underline{3}s^{1}$

➢ Period 4 (N)

- Highest principal quantum number **n=4**
- contains 18 elements beginning with potassium and ending with krypton

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

➢ Period 5 (O)

- Highest principal quantum number **n=5**
- contains 18 elements, beginning with rubidium and ending with xenon

Example: $_{38}Sr: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ \underline{5}s^2$

➢ Period 6 (P)

- Highest principal quantum number **n=6**
- contains 32 elements beginning with caesium and ending with radon

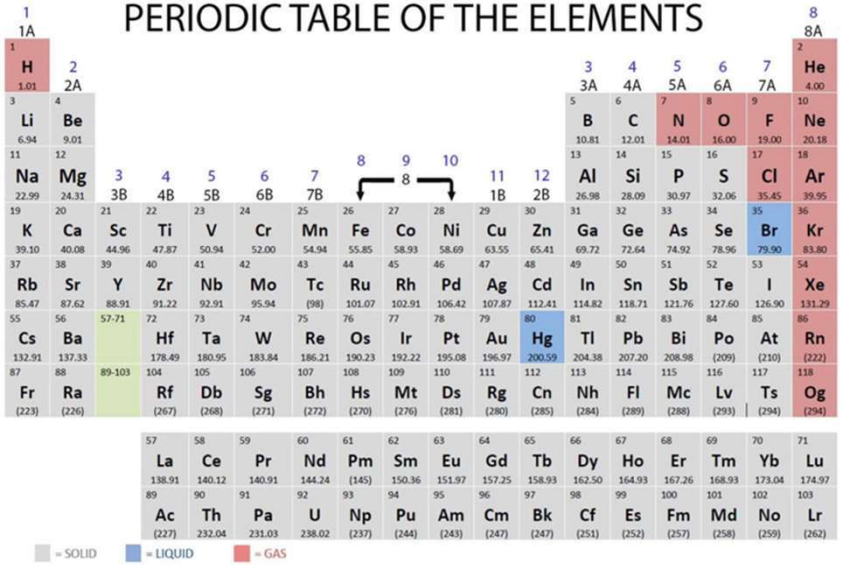
Example: $_{58}Ce: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ \underline{6}s^2\ 4f^2$

➢ Period 7 (Q)

- Highest principal quantum number **n=7**
- ontains 32 elements beginning with francium

Example: $_{87}Fr: [Rn]\ \underline{7}s^1$

3. Groups of the Periodic Table



- **The elements are arranged in 18 vertical columns called groups. Groups are labeled at the top of each column.**
- **The elements can also be classified into the main-group elements (or representative elements) in the columns labeled 1, 2 and 13–18 Group A and transition metals located in columns 3-12 Group B**

3.1. Group A elements

How to find the group for an element in the periodic table ?

1. write the electronic configuration of the element
2. Get the **total number** of electrons (e) in the outer **subshells/incomplete subshells (ns,np):** This gives the 'Group number'

➢ Group 1A

Elements of group 1A (**alkali metals**) have only one valence electron in their highest-energy orbitals (ns^1).

Example: $_3Li : 1s^2\ \mathbf{2s^1}$ group 1A

➢ Group 2A

Elements of group 2A (**alkaline earth metals**) have only one valence electron in their highest-energy orbitals (ns^2).

Example: $_4Be : 1s^2\ \mathbf{2s^2}$ group 2A

➢ Group 3A

Elements of group 3A have **three** valence electron in their highest-energy orbitals **(ns^2p^1).**

Example: $_5B : 1s^2$ **$2s^2\ 2p^1$** **group 3A**

➢ Group 4A

Elements of group 4A have **four** valence electron in their highest-energy orbitals **(ns^2p^2).**

Example: $_6C : 1s^2$ **$2s^2\ 2p^2$** **group 4A**

➢ Group 5A

Elements of group 5A have **five** valence electron in their highest-energy orbitals **(ns^2p^3).**

Example: $_7N : 1s^2$ **$2s^2\ 2p^3$** **group 5A**

➢ Group 6A

Elements of group 6A have **six** valence electron in their highest-energy orbitals **(ns^2p^4).**

Example: $_8O : 1s^2$ **$2s^2\ 2p^4$** **group 6A**

➢ Group 7A

Elements of group 7A (**halogens**) have **seven** valence electron in their highest-energy orbitals **(ns^2p^5).**

Example: $_9F : 1s^2$ **$2s^2\ 2p^5$** **group 7A**

➢ Group 8A

Elements of group 8A (**noble gases** or **inert gases**) have **six** valence electron in their highest-energy orbitals **(ns^2p^6).**

Example: $_{10}Ne : 1s^2$ **$2s^2\ 2p^6$** **group 8A**

3.2. Group B elements

- The **transition elements** or **transition metals** occupy the short columns in the center of the periodic table, between Group 2A and Group 3A. They are sometimes called the ***d*-block elements**,
- the transition metals in general have electron configurations of $ns^2\ (n-1)d$

How to find the group for Transition elements?

1. write the electronic configuration of the element

➢ **If the number of electrons (e) in *d*-orbitals is less than five**

Get the **total number** of electrons (e) in the outer **subshells ns, (n-1)d :** This gives the 'Group number'

Examples:

$_{21}Sc: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$	**Group 3B**
$_{22}Ti: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$	**Group 4B**
$_{23}V: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$	**Group 5B**
$_{24}Cr: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$	**Group 6B**
$_{25}Mn: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$	**Group 7B**

➢ **If the number of electrons (e) in *d*-orbitals is more than five and less than ten**

The group is 8B

Examples:

$_{26}$Fe: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^2\ 3d^6}$	**Group 8B**
$_{27}$Co: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^2\ 3d^7}$	**Group 8B**
$_{28}$ Ni: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^2\ 3d^8}$	**Group 8B**

➢ **If the number of electrons (e) in *d*-orbitals equals ten**

Get the **total number** of electrons (e) in the outer **subshells s:** This gives the 'Group number'

Examples:

$_{29}$Cu: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^1\ 3d^{10}}$	**Group 1B**
$_{27}$Zn: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^2\ 3d^{10}}$	**Group 2B**

Examples:

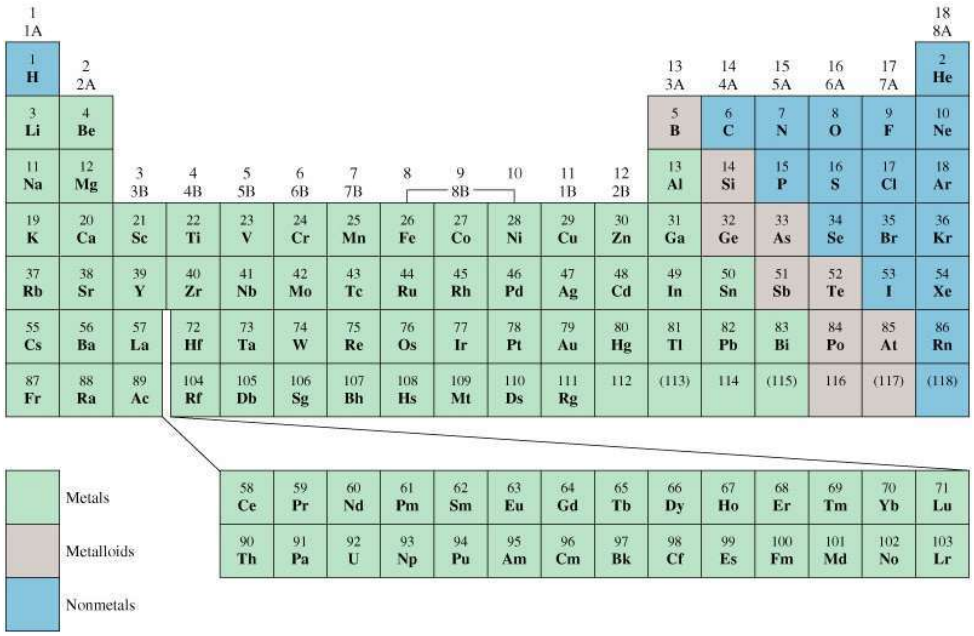
1- The element with the valance electron configuration of $3s^2\ 3p^3$ is:

a) Carbon

b) Nitrogen

c) Phosphorus

d) Neon



2- Which of the following sets of elements are expected to have similar chemical properties?

a) Sulfur and phosphorous

b) Sulfur and oxygen

c) Sulfur and argon

3-Titanium (Ti) element is found in the periodic table in:

a) s-block
b) P-block
c) d-block
d) f-block

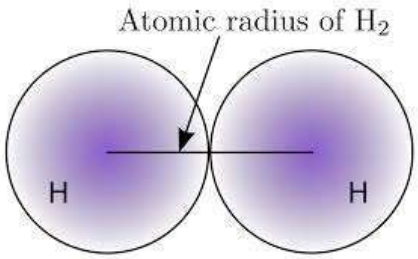
4- The general formula of an element in group 8A is

a) S^2
b) S^1
c) s^2p^6
d) s^1p^1

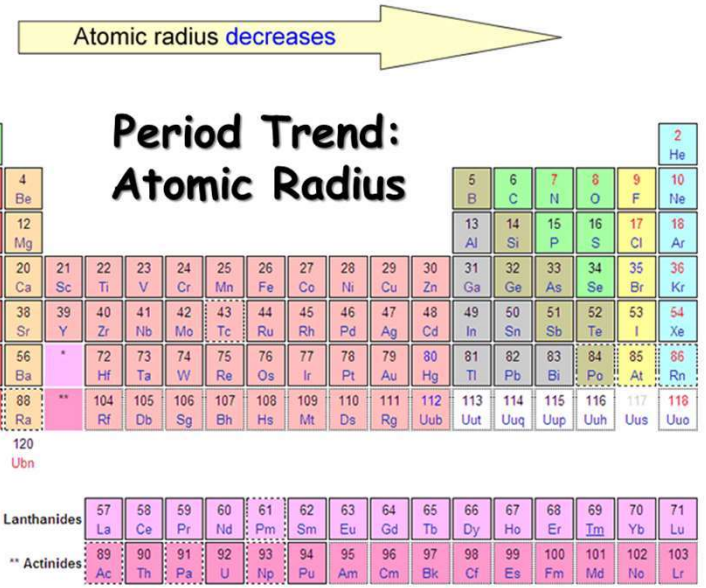
5- Iron is located in group --------- and period ----------

4. Periodic Variation in Physical Properties

4.1. Atomic Radius



- The size of the atom is defined by the ATOMIC RADIUS.
- **Atomic Radius (AR) :** is one-half the distance between the two nuclei in two adjacent metal atoms or in a diatomic molecule.
- **In the same period:** n is the same → the added valence electron is more strongly attracted by the nucleus → As a result, the atomic radius decreases from left to right across the period.



Examples:

Element	period	group
$_4$Be: $1s^2 2s^2$	2	**2A**
$_5$B: $1s^2 2s^2 2p^1$	2	**3A**
$_6$C: $1s^2 2s^2 2p^2$	2	**4A**

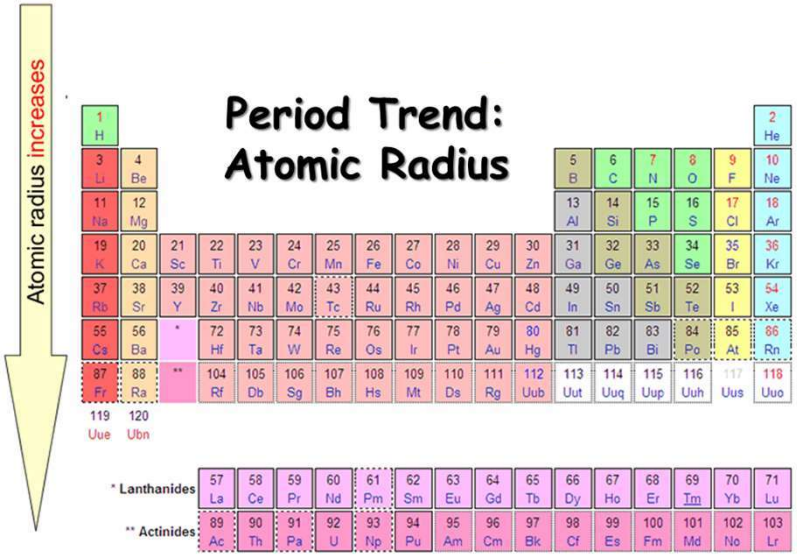
the atomic radius decreases from Beryllium (Be) to Carbon (C) (same period)

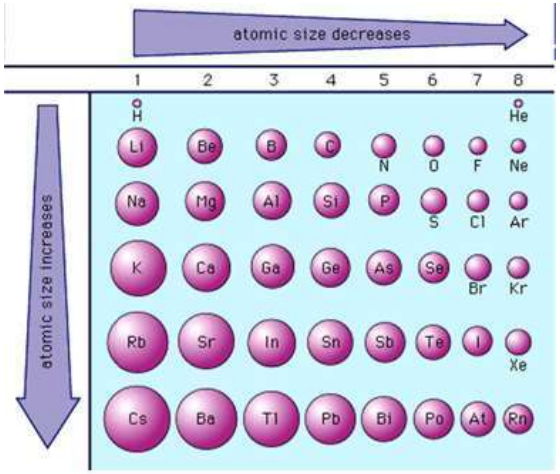
In the same group: The atomic radius increases from top to bottom as the electrons are added to new energy shells (n increases).

Examples:

Element	period	group
$_{1}$H: **1**s^{1}	1	**1A**
$_{3}$Li: **1**s^{2}**2**s^{1}	2	**1A**
$_{11}$Na :1s^{2} 2s^{2} 2p^{6} **3**s^{1}	3	**1A**

the atomic radius increases from Hydrogen (H) to Sodium (Na) (same group)





Referring to a periodic table, arrange the following atoms in order of increasing atomic radius: P, Si, N.

Increasing means from small to large

(Small) N < P < Si (Large)

Arrange the following atoms in order of decreasing radius: C , Li, Be .

Decreasing means from large to small

(Large) Li > Be > C (Small)

Which choice below correctly lists the elements in order of increasing atomic radii?

a- Na < Mg < K < Rb

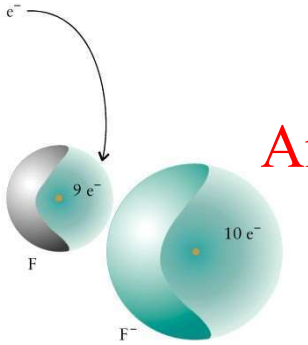
b- Mg < Na < K < Rb

c- Rb < K < Na < Mg

d- Rb < K < Mg < Na

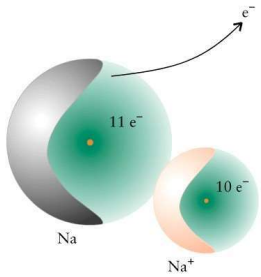
4.2. Ionic Radius

- **Ionic Radius:** is the radius of cations or anions.
- **Anions:** gain electrons → The **ionic radius increases** because the nuclear charge remains the same but the repulsion resulting from the additional electrons increases the size.



Anion

- **Cations:** lose electrons → The **ionic radius decreases** because the removal of one or more electrons from an atom reduces electron-electron repulsion while the nuclear charge remains the same. This leads to shrink the electron clouds.

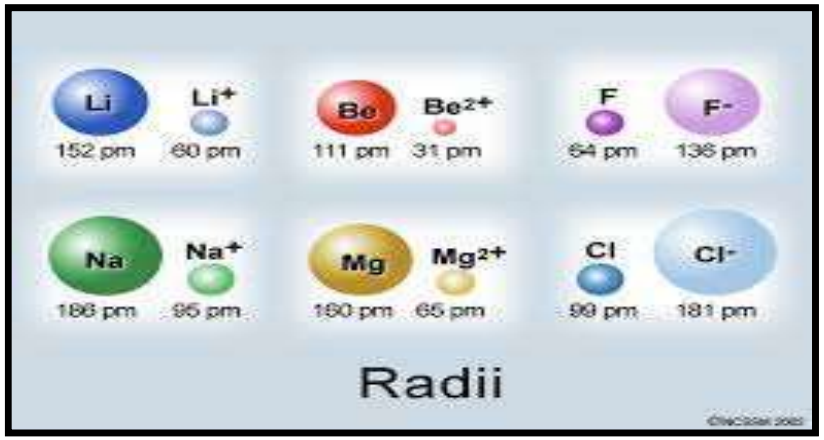


Cation

Na> Na^+

$_{11}Na$: $1s^2\ 2s^2\ 2p^6\ \mathbf{3}s^1$

$_{11}Na^+$: $1s^2\ 2s^2\ 2p^6$



F <F^-

$_9F$: $1s^2\ 2s^2\ 2p^5$

F^- : $1s^2\ 2s^2\ 2p^6$

Order the following according to the increase in the atomic/ionic radius.

N^{3-} Li^{+} C O^{2-}

a) $C < Li^{+} < O^{2-} < N^{3-}$

b) $N^{3-} < O^{2-} < C < Li^{+}$

c) $Li^{+} < C < N^{3-} < O^{2-}$

d) $Li^{+} < C < O^{2-} < N^{3-}$

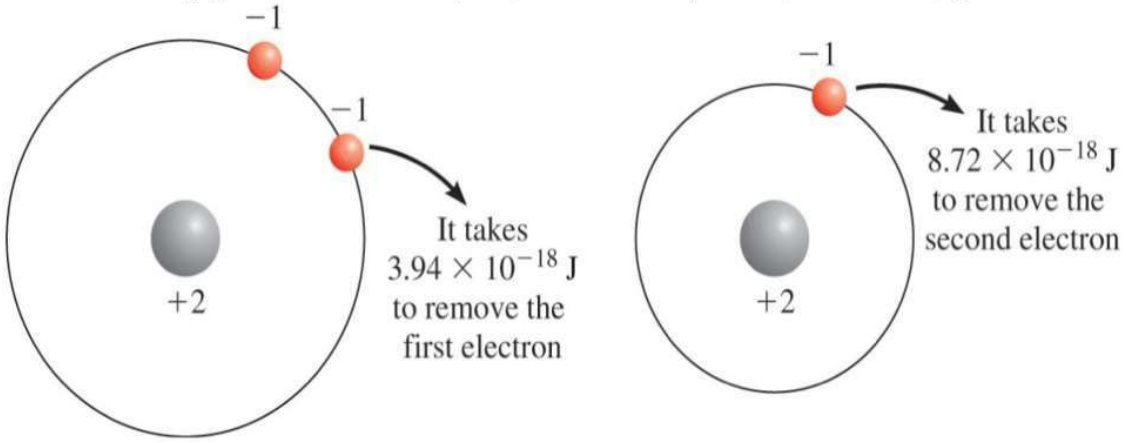
4.3. Ionization Energy

- **Ionization Energy (IE) :** is the minimum energy in (kJ/mol) required to remove an electron from a gaseous atom in its ground state.
- For a many-electron atom, the amount of energy required to remove the first electron from the atom in its ground state is called the first ionization energy (I_1). The amounts of energy required to remove the second and third electrons are called the second and third ionization energies (I_2 and I_3), respectively.

- When an electron is removed from the atom, the repulsion among the remaining electrons decreases because the nuclear charge remains constant. However, more energy is needed to remove another electron from the positively charged ion.

$$I_1 < I_2 < I_3$$

Copyright © The McGraw-Hill Companies, Inc. Permission required for reproduction or display.

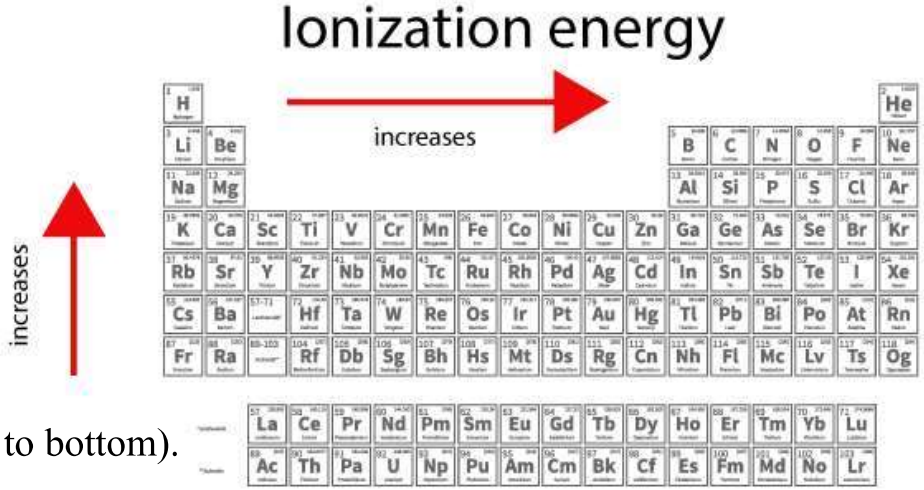


❖ **In the same period:** n is the same → the added valence electron is more strongly attracted by the nucleus → As a result, the ionization energy increases from left to right across the period.

Examples:

Element	period	group
$_5$B: $1s^2\mathbf{2}s^2 2p^1$	2	**3A**
$_6$C: $1s^2\mathbf{2}s^2 2p^2$	2	**4A**
$_7$N: $1s^2\mathbf{2}s^2 2p^3$	2	**5A**

IE: $_7$N > $_6$C > $_5$B



❖ **In the same group:** The IE decreases down the groups (from top to bottom).

Examples:

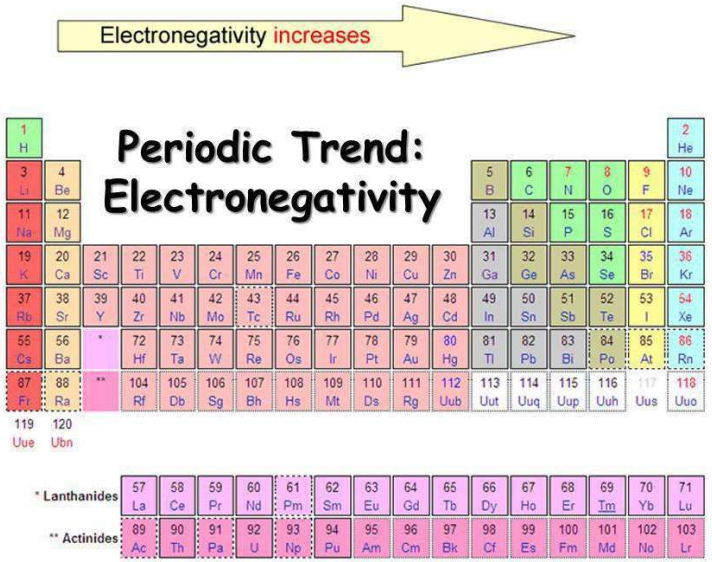
Element	period	group
$_3$Li: $1s^2\mathbf{2}s^1$	2	**1A**
$_{11}$Na :$1s^2\ 2s^2\ 2p^6\ \mathbf{3}s^1$	3	**1A**
$_{19}$K :$1s^2\ 2s^2\ 2p^6\ \mathbf{3}s^2 3p^6\ \mathbf{4}s^1$	4	**1A**

IE: $_3$Li > $_{11}$Na > $_{19}$K

4.4. Electronegativity

- **Electronegativity:** is the ability of an atom to attract toward itself the electrons in a chemical bond.
- High electronegativity → able to pick up electrons easily
- Nonmetals have high electronegativity, On the other hand, metals have lower electronegativity.

❖ **In the same period:** n is the same → the added valence electron is more strongly attracted by the nucleus → As a result, the Electronegativity increases from left to right across the period.



Examples:

Element	period	group
$_5$B: $1s^2 2s^2 2p^1$	2	**3A**
$_6$C: $1s^2 2s^2 2p^2$	2	**4A**
$_7$N: $1s^2 2s^2 2p^3$	2	**5A**

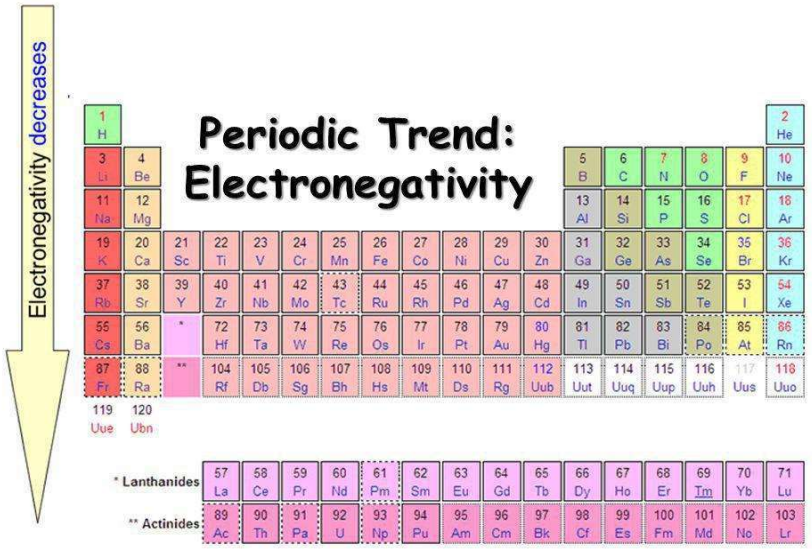
Electronegativity: $_7N > {}_6C > {}_5B$

❖ **In the same group:** The electronegativity decreases down the groups (from top to bottom).

Examples:

Element	period	group
$_{3}$Li: **1**s^2**2**s^1	2	**1A**
$_{11}$Na :1s^2 2s^2 2p^6 **3**s^1	3	**1A**
$_{19}$K :1s^2 2s^2 2p^6 **3**s^{2}3p^6 **4**s^1	4	**1A**

Electronegativity: $_{3}$Li > $_{11}$Na > $_{19}$K



Example:

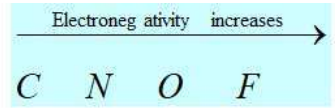
Which choice below correctly lists the elements in order of increasing electronegativity?

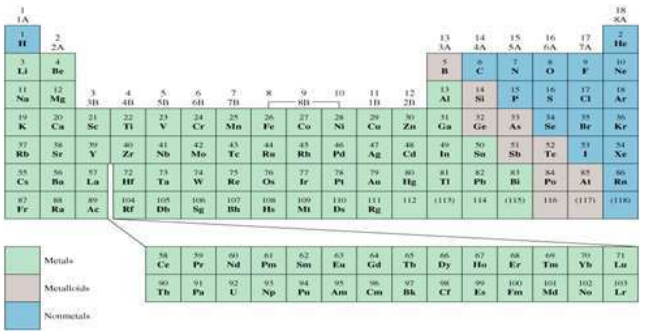
a) **C < N < O < F**

b) N < C < O < F

c) N < C < F < O

d) C < N < F < O





Which atom should have a smaller first ionization energy: oxygen or sulphur?

Both oxygen and sulphur atoms are in group 6 but different period, the valence electrons in sulphur are farther from the nucleus → removing of them is easier

THUS: I_1 (S) < I_1 (O)

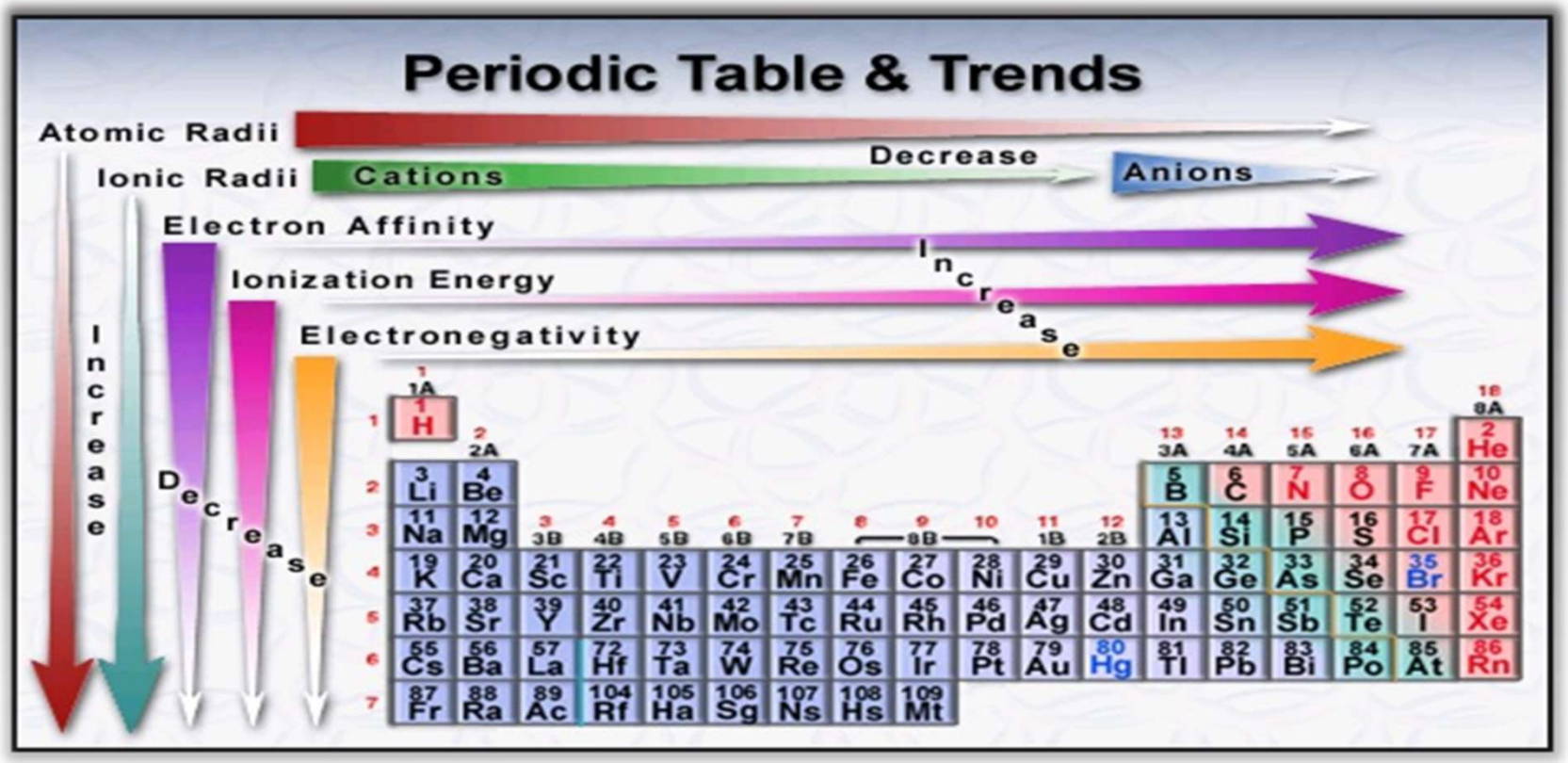
Which atom should have a higher second ionization energy (Li or Be)?

Li ($2s^1$) < Be ($2s^2$) … first ionization energy

Li^+ ($1s^2$) > Be^+ ($2s^1$) … second ionization energy

Because 1s electrons shield 2s electrons much more effectively than they shield each other , we predict that it should be easier to remove a 2s electron from Be^+ than to remove a 1s electron from Li^+

Summary


Periodic Table & Trends
Atomic Radii
Ionic Radii
Cations
Decrease
Anions
Electron Affinity
Ionization Energy
Increase
Electronegativity
Increase
Decrease