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CHAPTER 3.0
PERIODIC TABLE

CHEMISTRY 1
SK015

SESSION 2025/2026

STUDENT LEARNING TIME (SLT): LECTURE	
NON FACE-TO-FACE (PREPARATION)	FACE-TO-FACE (DURING CLASS)
2 HOURS	2 HOURS

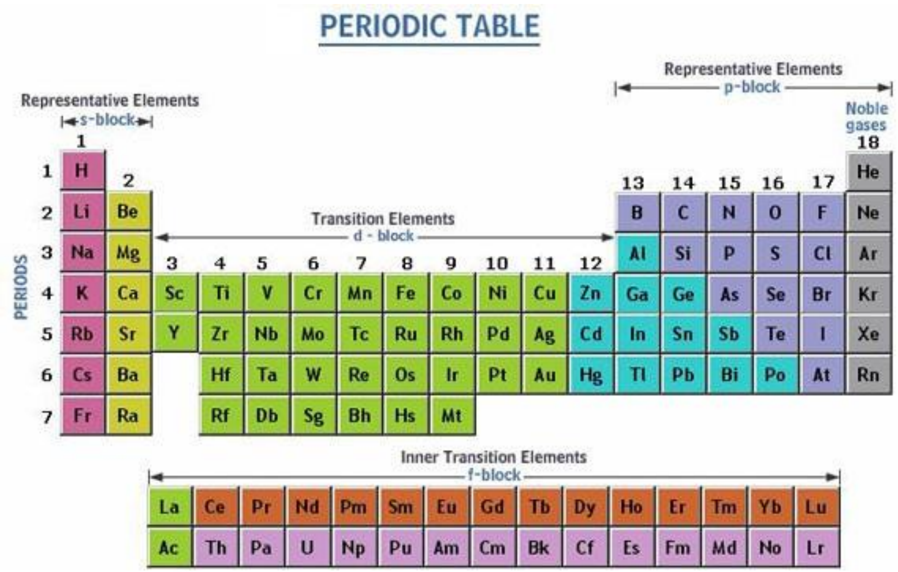
Chemistry unit

CHEMISTRY UNIT, KMJ

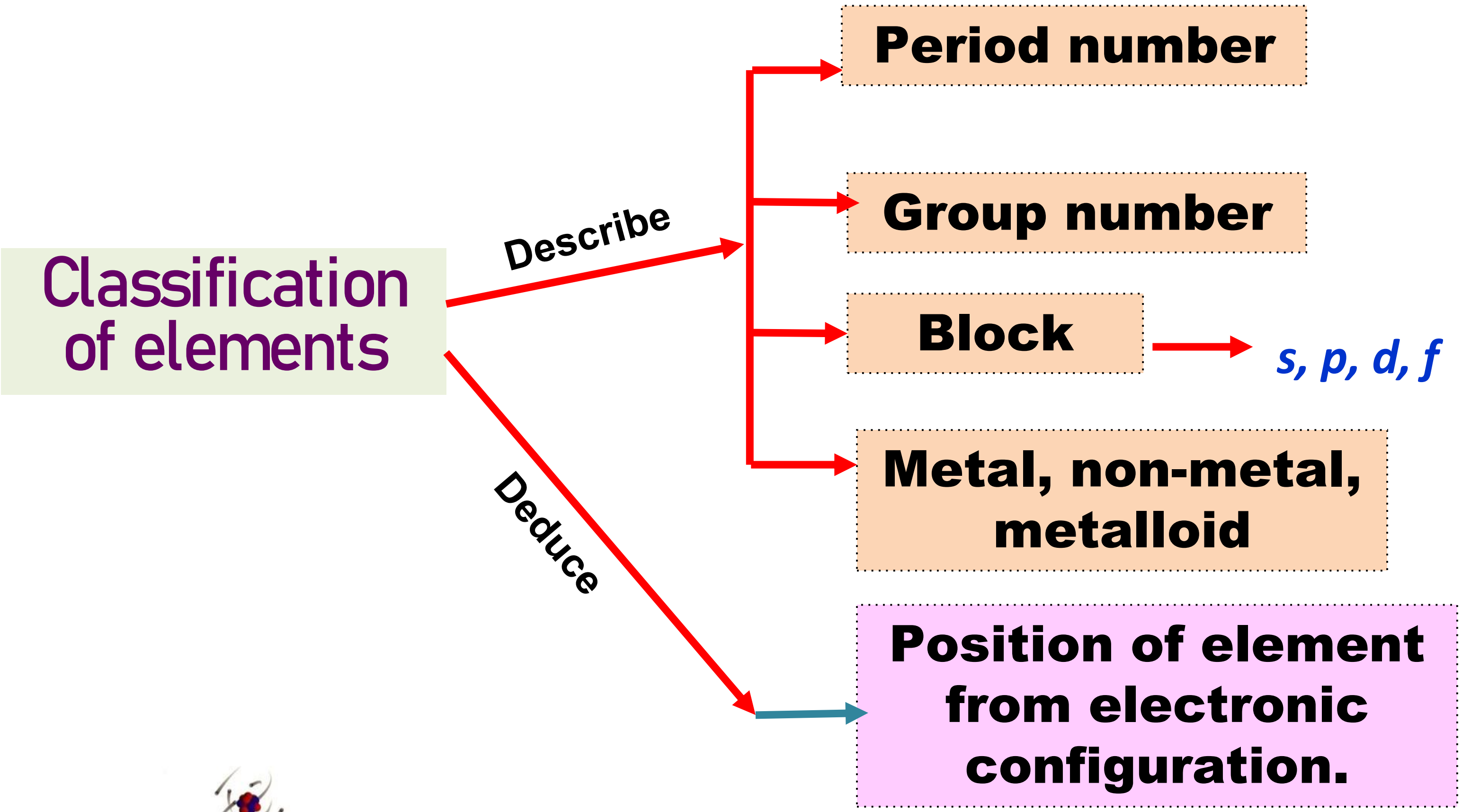
3.0
PERIODIC TABLE

3.1
CLASSIFICATION OF ELEMENTS

3.2
PERIODICITY



CHAPTER 3.1 : OVERVIEW

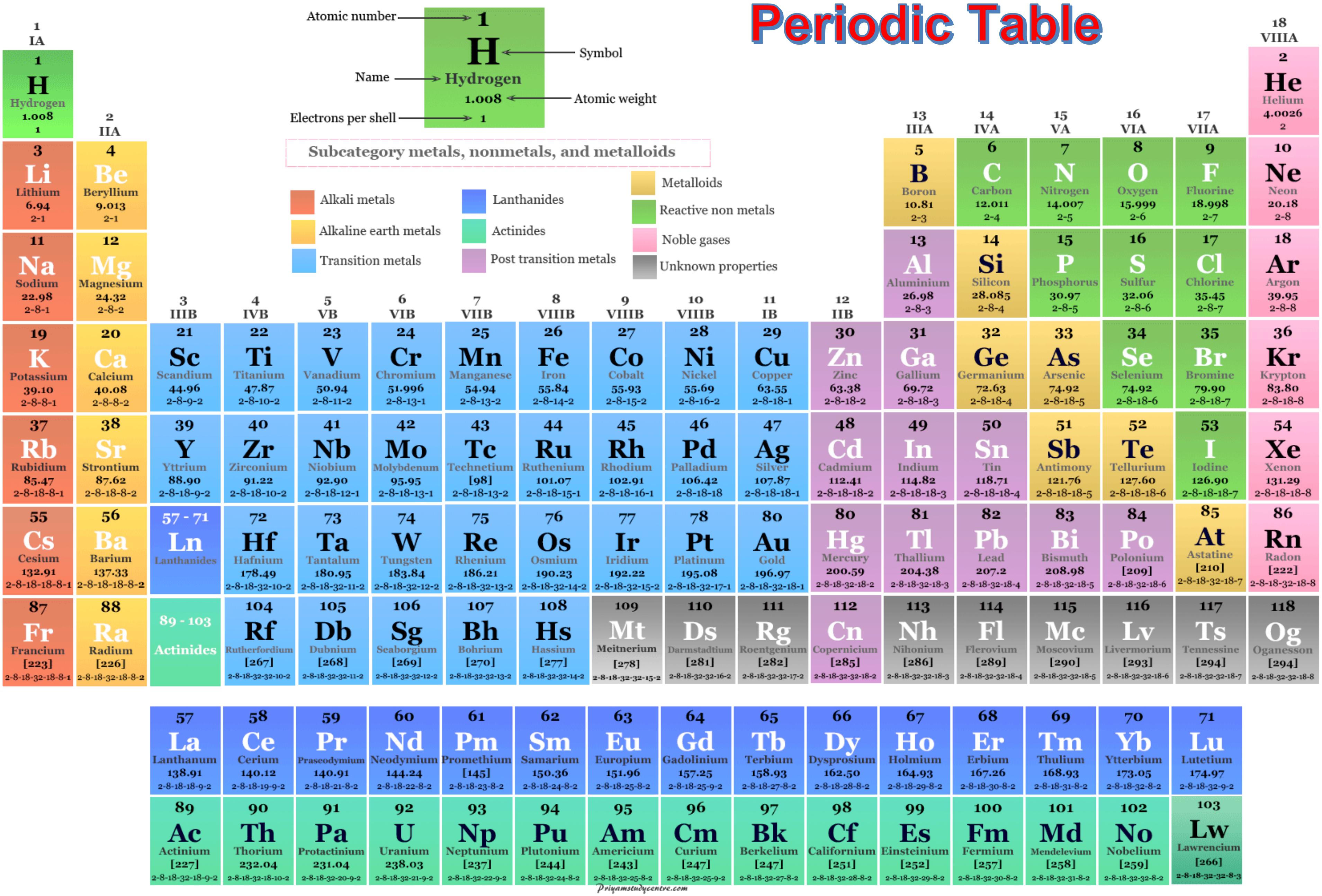


3.1 CLASSIFICATION OF ELEMENTS

Teaching and learning outcomes

At the end of the lesson, student should be able to

3.1	**Classification of elements**
a)	Describe period, group, and block (*s,p,d,f*) (C1, C2)
b)	Deduce the position of elements in the periodic table from its electronics configuration (C4)



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CLASSIFICATION OF ELEMENTS

- In periodic table, elements are arranged in order of increasing **proton number (Z)**.
- A **vertical column of elements** is called a **group.**
- A **horizontal row** is known as a **period**.

GROUPS

- The group in the Periodic Table are numbered from **1 to 18**.

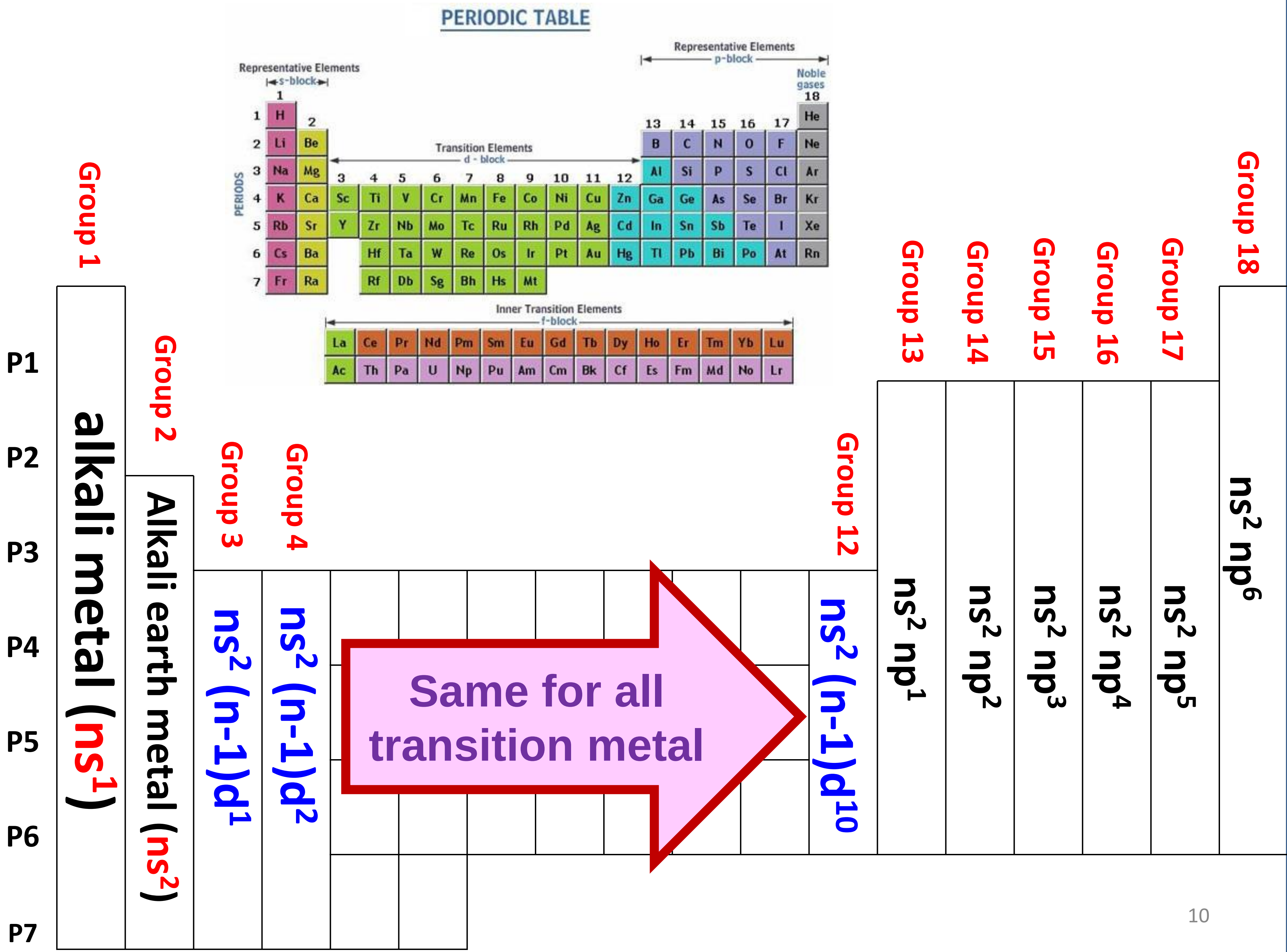
- Elements in the **same group** have the **same number of valence electron**.

- Example:

  Sodium(Na) and potassium(K) are both found in **group 1** which means that they both have **1 valence electron**.

- **Main group (Representative Group) in Periodic Table and indicates valence electron :**

GROUP	Valence electron configuration
Group 1 (alkali metals)	ns^1
Group 2 (alkali earth metals)	ns^2
Group 3 – 12 (transition metals)	$ns^2(n-1)d^1$ to $ns^2(n-1)d^{10}$
Group 13	$ns^2\ np^1$
Group 14	$ns^2\ np^2$
Group 15	$ns^2\ np^3$
Group 16 (chalcogen)	$ns^2\ np^4$
Group 17 (halogen)	$ns^2\ np^5$
Group 18 (inert/ noble gases)	$ns^2\ np^6$

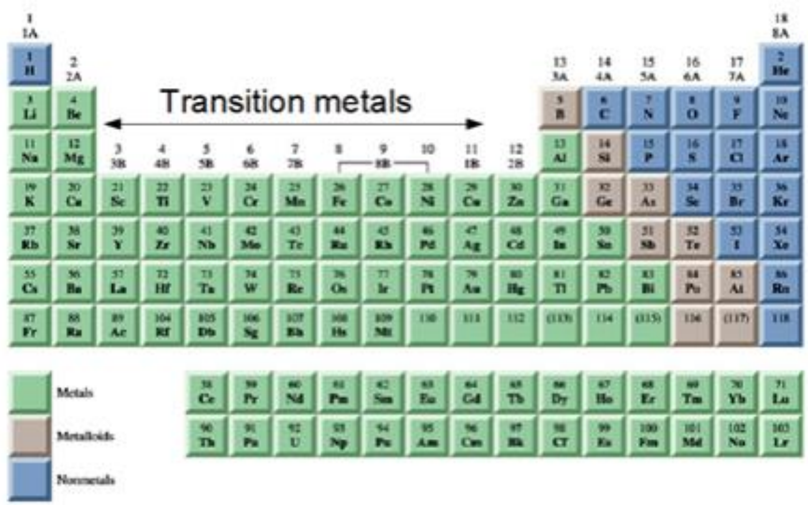

PERIODIC TABLE
Representative Elements
s-block
Representative Elements
p-block
Noble gases
Transition Elements
d - block
Inner Transition Elements
f-block
PERIODS
Group 1
Group 2
Group 3
Group 4
Group 12
Group 13
Group 14
Group 15
Group 16
Group 17
Group 18
P1
P2
P3
P4
P5
P6
P7
alkali metal (ns1)
Alkali earth metal (ns2)
ns2 (n-1)d1
ns2 (n-1)d2
Same for all transition metal
ns2 (n-1)d10
ns2 np1
ns2 np2
ns2 np3
ns2 np4
ns2 np5
ns2 np6

PERIOD

- **Period are numbered 1 to 7**

- ***How to determine period ?***

  It can be determined by referring to the valence electrons and its principal quantum number, n



EXAMPLE:

Sodium ($_{11}Na$)= $1s^2\ 2s^2\ 2p^6$ <u>$3s^1$</u>

Sodium is an element in period 3 because the valence electron is filled the outermost shell (n = 3)

EXAMPLE:

Chlorine($_{11}$Cl)= $1s^2\ 2s^2\ 2p^6$ <u>$3s^2\ 3p^5$</u>

Chlorine is an element in period 3 because the valence electron is filled the outermost shell (n = 3)

BLOCKS

- All the elements in the Periodic Table can be classified into **4 main blocks** according to their valence electrons configuration.
- These main blocks are:

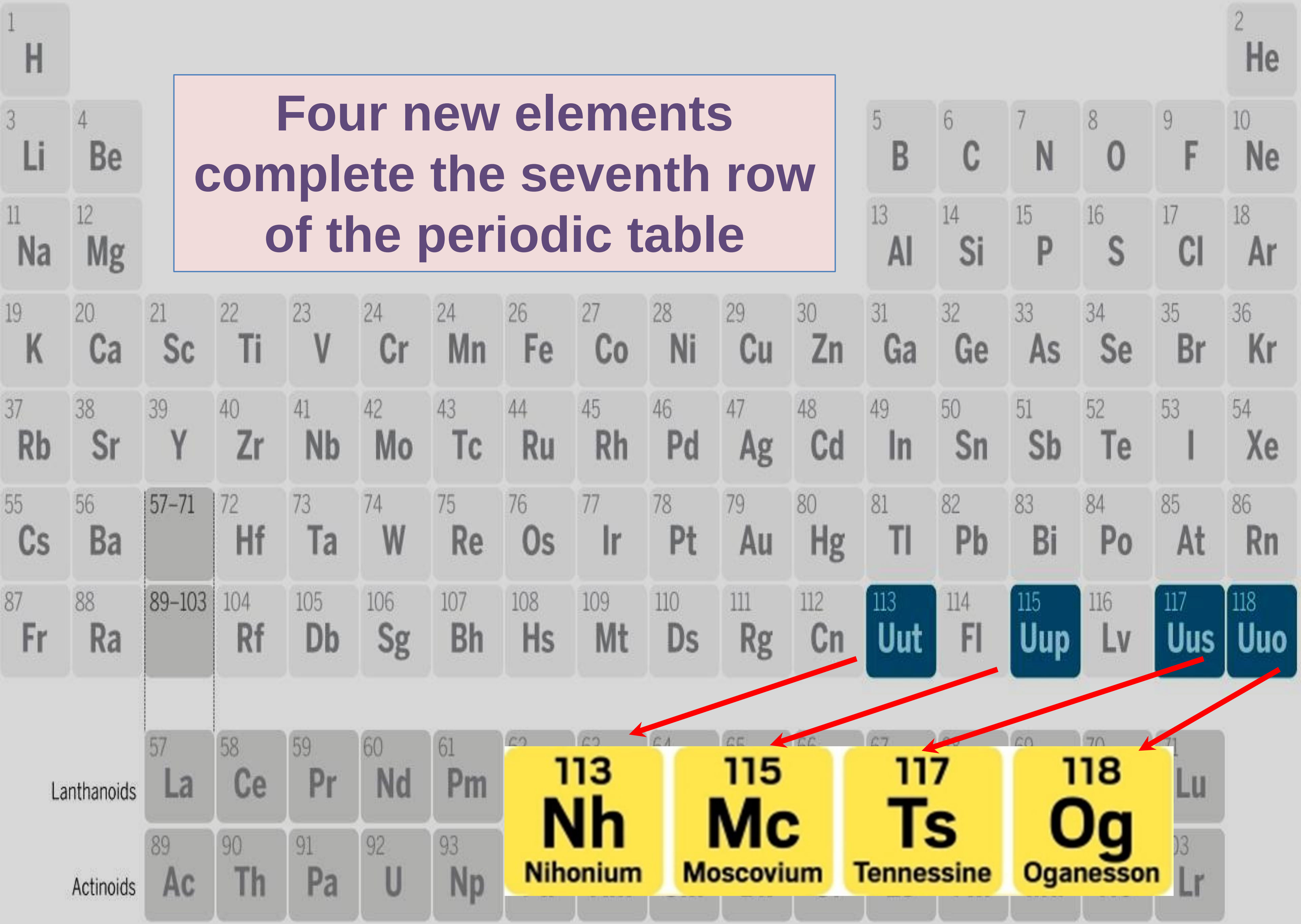
i. *s* block (group 1 & 2)

ii. *p* block (group 13 - 18)

iii. *d* block (group 3 - 12)

iv. *f* block (lanthanides and actinides series)

Four new elements complete the seventh row of the periodic table



s - BLOCK

Group 1 and 2

- The filling of valence electrons involve the *s* orbital
- Configuration of the valence electrons : ns^1 to ns^2

EXAMPLE:

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

$_{20}Ca$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

p - BLOCK

Group 13 to 18

- **The filling of valence electrons involve s and p orbitals**
- **The configuration of valence electrons: $ns^2\ np^1$ to $ns^2\ np^6$**

EXAMPLE:

$_{13}Al$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

$_{52}Te$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^4$

d - BLOCK

Group 3 to 12

- **Also known as a transition elements.**
- **The filling of valence electrons involve ns and (*n-1*)*d* orbital**
- **Configuration of valence electron :**
  ***ns*^{2} (*n-1*)d^1 to *ns*^{2}(*n-1*)*d*^{10}**

EXAMPLE:

$_{23}V$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$

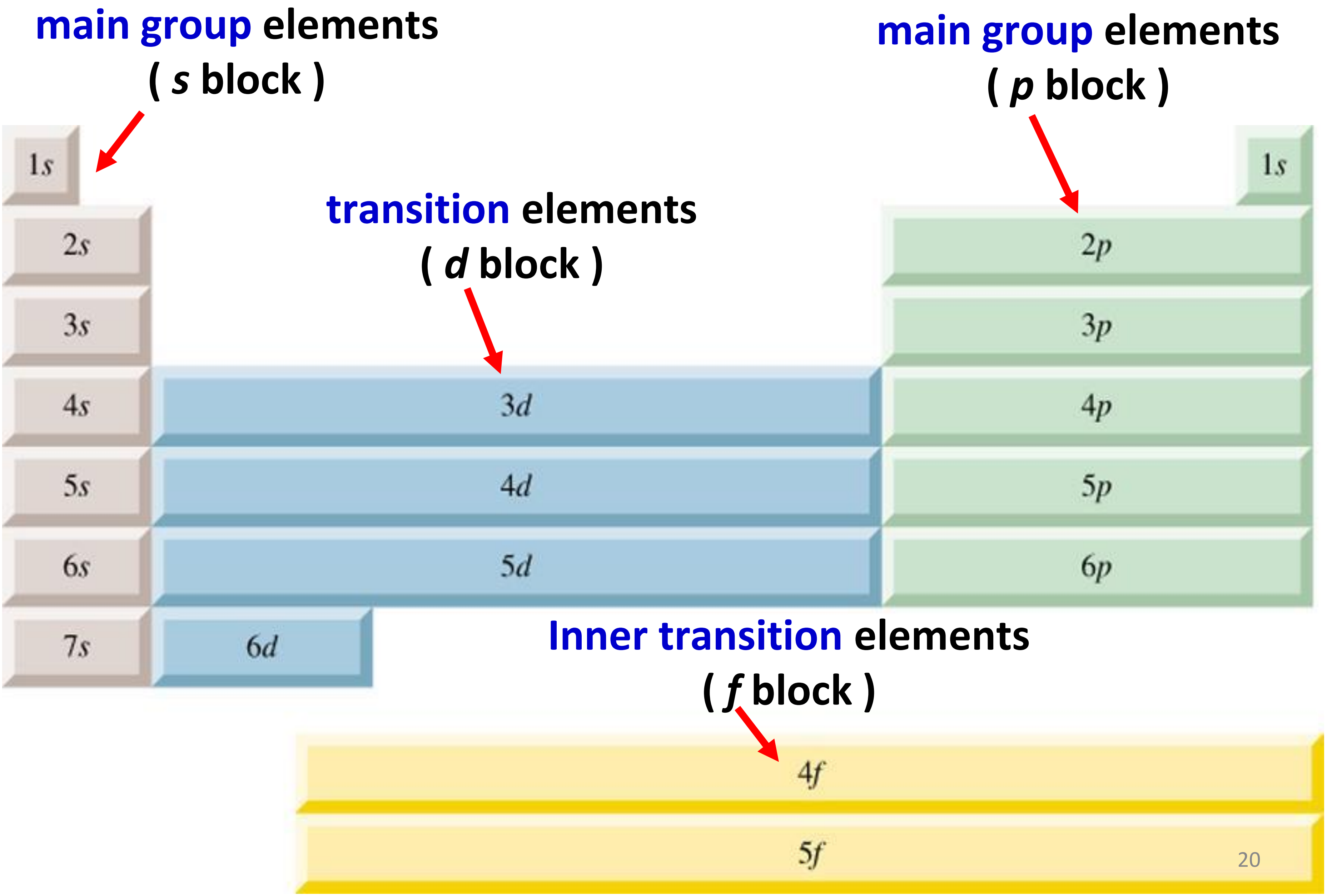
or

[Ar] $3d^3\ 4s^2$ where [Ar] = 18 electrons

f - BLOCK

- **Involve the elements in the series of lanthanides (Ce to Lu) and actinides (Th to Lr)**

- **The filling of valence electrons happen in the subshell of 4f and 5f**


main group elements
(s block)
main group elements
(p block)
transition elements
(d block)
Inner transition elements
(f block)
1s
2s
3s
4s
5s
6s
7s
1s
2p
3p
4p
5p
6p
3d
4d
5d
6d
4f
5f

EXAMPLE:

Without referring to the periodic table, classify these elements according to their group, period and block .

(a) K (Z = 19)
(b) Cl (Z = 17)
(c) Ti (Z = 22)

Answer:

(a) K (Z = 19)

Electron configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

Number of valence electron = 1

Group 1 Period 4 Block *s*

(b) Cl (Z = 17)

Electron configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Number of valence electron = 7

Group 17 Period 3 Block *p*

Answer:

(c) Ti (Z = 22)

Electron configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

Number of valence electron = 4

Group 4 Period 4 Block *d*

METALS, NON-METALS, METALLOIDS

- **Metals** are located in the **left side** and **in the middle** of the periodic table

- **Nonmetals** are located in the **upper right quarter** of the table

- **Metalloids (semi-metal)** are elements that lie along the line that separates metal from non-metal.

3.2 PERIODICITY

Teaching and learning outcomes

At the end of the lesson, student should be able to

3.2	**Periodicity**
a)	Describe the variation in atomic radii : (C1, C2) (i) across period (ii) across the first row of transition element (iii) down a group
b)	Analyse the variation in atomic radii : (C4) (i) across period (ii) across the first row of transition element (iii) down a group
c)	Compare the atomic radius of an element and its corresponding ionic radius (C2, C4)

3.2 PERIODICITY

Teaching and learning outcomes

At the end of the lesson, student should be able to

3.2	Periodicity
d)	Define the terms isoelectronic (C1)
e)	Compare the radius of isoelectronic species (C2, C4)
(f)	Analyse the variation in the ionic radii across period 2 and period 3 (C4)
g)	Define the first and second ionization energy (C1)
h)	Analyse the variations in the first ionisation energy across period and down a group (C4)
i)	Explain the increase in the successive ionization energies of an element (C2, C3)

3.2 PERIODICITY

Teaching and learning outcomes

At the end of the lesson, student should be able to

3.2	**Periodicity**
j)	Deduce the electronic configuration of an element and its position in the periodic table based on successive ionisation energy data (C4)
k)	Define electronegativity (C1)
l)	Explain the variation in electronegativity of elements (C2, C3)
m)	Explain the acids base character of oxides of elements in period 3 (C2, C3) **relate the acid-base character to the types of bonding (ionic or covalent)*

NUCLEAR CHARGE , Z

- **Number of protons $\uparrow$, $Z \uparrow$**

- **Nucleus–electron attraction become stronger**

EFFECTIVE NUCLEAR CHARGE

- is the **charge** felt by the valence electrons as a results of **shielding effect** by other electrons

- Abbreviation : Z_{eff}

- When Z_{eff} increases, the nucleus attraction towards electrons become stronger

FACTORS AFFECTING ATOMIC SIZE

1) Change in effective nuclear charge, Z_{eff}

- As Z_{eff} increase
- The positive charge "felt" by an e^- increases
- Outer e^- are pulled closer to nucleus
  ☞ nucleus–electron attraction increases

- Atoms become smaller

EXAMPLE:

Sodium, Na(Z = 11) ☞ Group 1

Full e^- configuration: $1s^2\ 2s^2\ 2p^6\ 3s^1$

Number of inner e^- = 10 **Z_{eff} = +1**

Magnesium, Mg= 12 ☞ Group 2

Full e^- configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2$

Number of inner e^- = 10 **Z_{eff} = +2**

So, magnesium smaller than sodium

2) Change in principle quantum number, *n* or sheilding effect

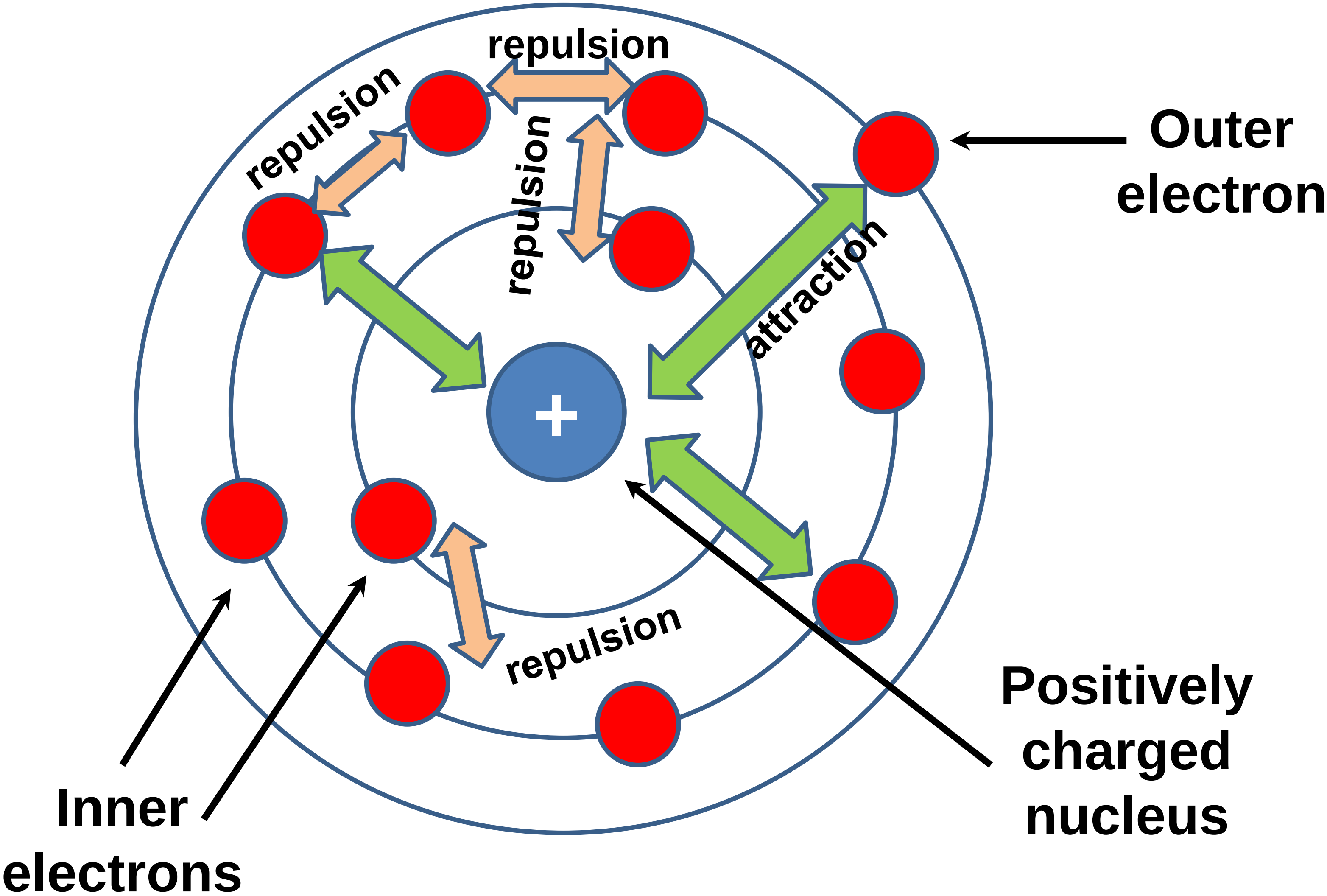
- As *n* increase, shielding effect also increase
- Outer e^- spend more time farther from nucleus
- Outer e^- are less attracted to the nucleus
  ☞ nucleus–electron attraction decreases
- Atoms are larger

SHIELDING EFFECT

- **Ability of other e^-, especially inner e^-, to lessen the nuclear attraction on an outer e^-**

- **Shielding effect occur between electrons of inner orbitals and the electrons occupying the valence orbitals.**

- **Some shielding effect also occur between electrons of the same orbital.**

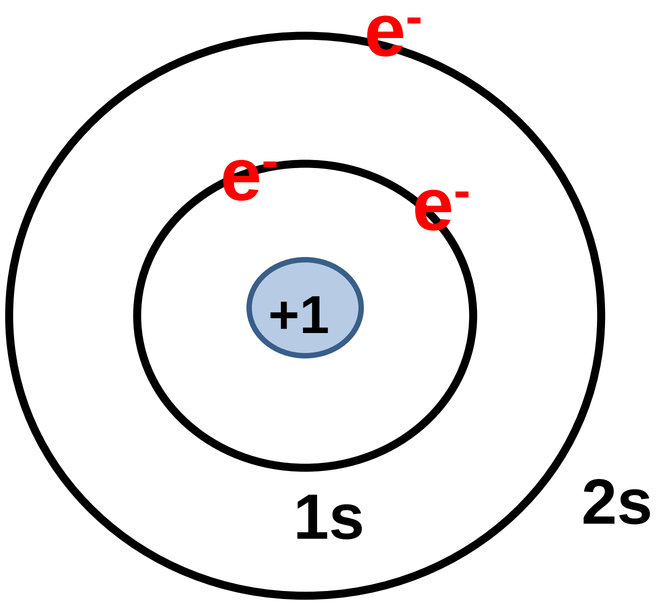


- Inner e^- shield outer e^- more effectively
- Outer electron felt less attraction from nucleus

EXAMPLE:

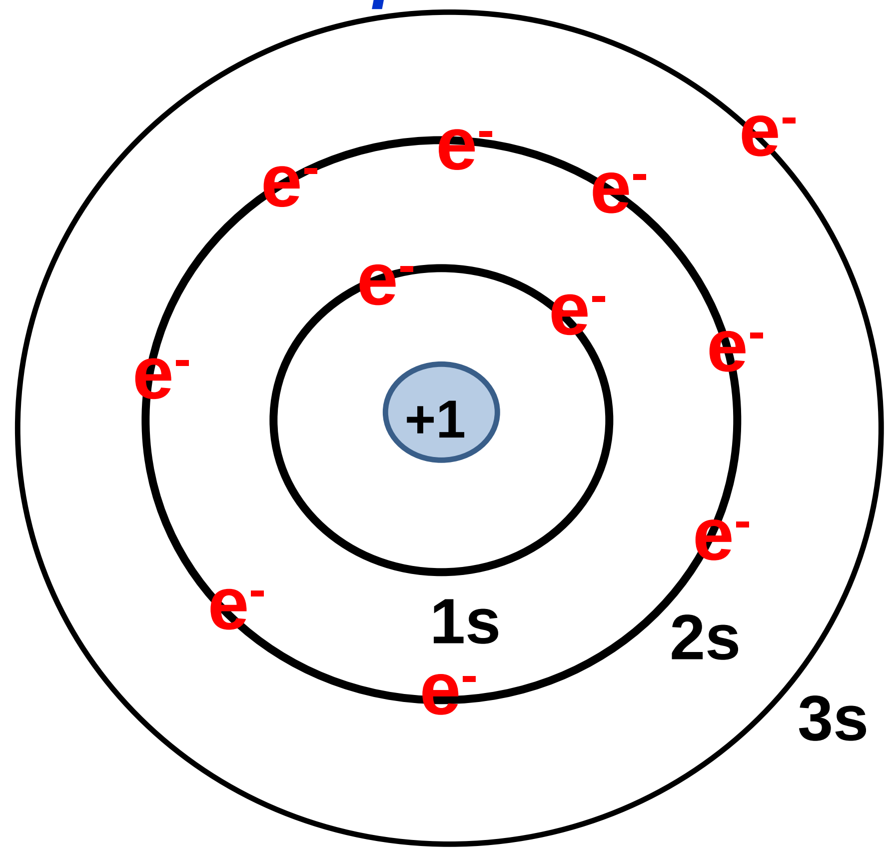
Li (Z = 19)

☞ **$1s^2$ $2s^1$**



Na(Z = 11)

☞ **$1s^2$ $2s^2$ $2p^6$ $3s^1$**



The outer 2s electron in Li atom is shielded effectively from the nucleus by the two 1s electrons.
The outer 3s electron in Na atom is shielded effectively from the nucleus by the 2 electrons in 1s and 8 electrons in 2s & 2p.

TRENDS IN ATOMIC SIZE

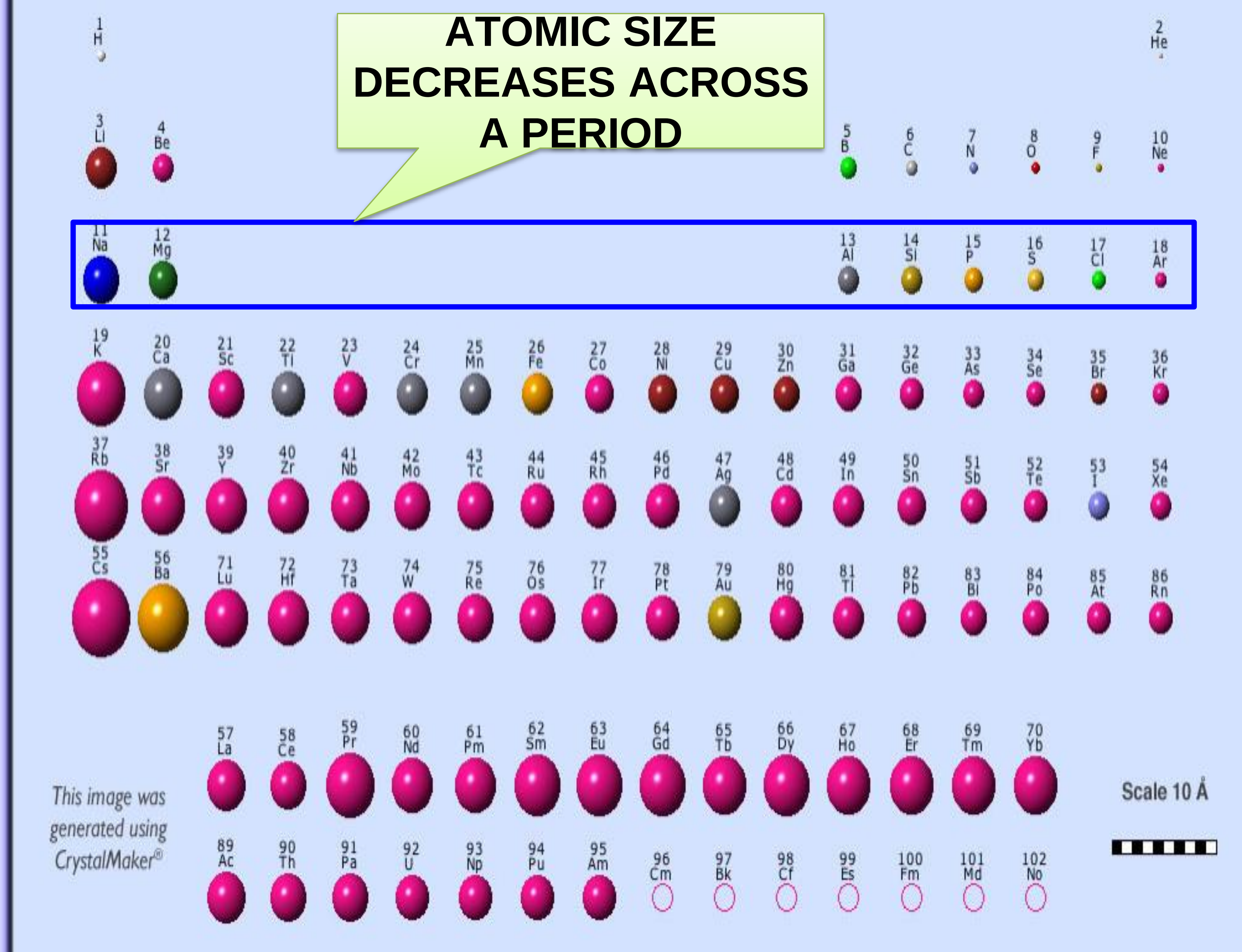
- **Atomic radius generally:**

  (i) Across a period from left to right , atomic radius decrease

  (ii) Down a group from top to bottom , atomic radius increase

Example :

Moving from Na to Al (period 3), the atomic size decreases


ATOMIC SIZE DECREASES ACROSS A PERIOD
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
71 Lu
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
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
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
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Scale 10 Å

(i) Across a Period , from left to right

- **The atomic radius of elements decreases**

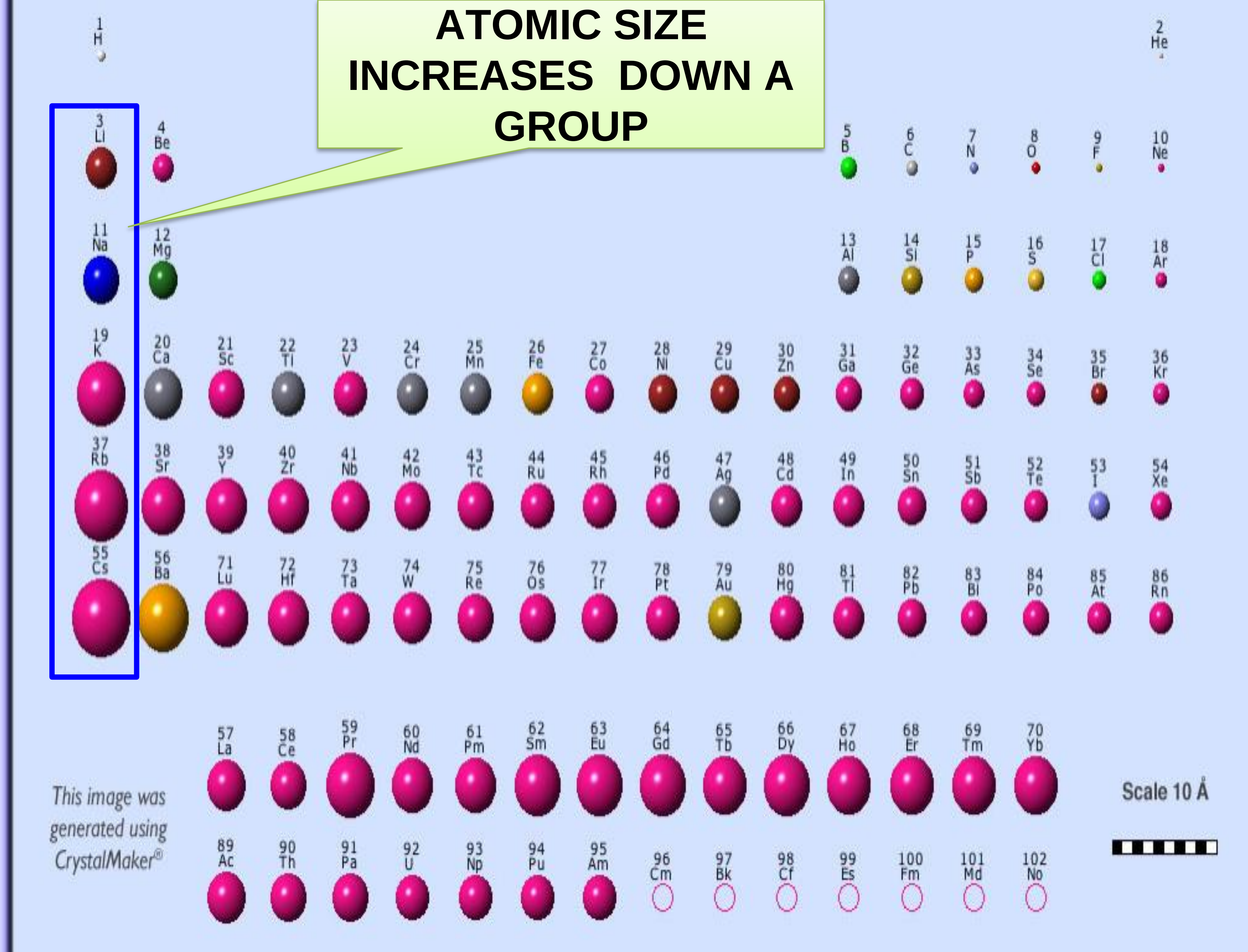
Reason:

- **Across the period, the proton number increases**
- **Effective nuclear charge increases**
- **Nucleus attraction towards valence (outer) electron increases**
- **Electrons are pulled closer to the nucleus**
- **Thus, atomic radii decreases**

Example :

As we move down from Li to K (Group 1), the atomic size increases

ATOMIC SIZE INCREASES DOWN A GROUP



(ii) Down a Group , from top to bottom

- The atomic radius of elements increases

Reason :

- Going down a group, the number of shell (*n*) increase due to the number of electrons increases
- Shielding effect increases
- The distance between valence electrons and nucleus are farther
- Nucleus attraction towards valence (outer) electrons weaker
- Thus, atomic radii increases

Example 2

By referring to the electronic configuration, arrange each set in order of increasing atomic size. Explain.

(a) $_{37}Rb$, $_{19}K$, $_{55}Cs$

(b) $_{12}Mg$, $_{19}K$, $_{20}Ca$

Answer

(a) $_{37}Rb$, $_{19}K$, $_{55}Cs$

$_{37}Rb$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$

period 5

$_{19}K$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

period 4

$_{55}Cs$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^1$

period 6

K < Rb < Cs

- K, Rb and Cs are in the same group. Going down a group, the number of shell (n) increases due to the number of electrons increases. The shielding effect also increases

- The distance between valence electrons and nucleus are farther.

- Nucleus attraction towards valence (outer) electron weaker

- Thus, atomic size increases from K to Cs.

Answer

(b) $_{12}Mg$, $_{19}K$, $_{20}Ca$

$_{12}Mg$: $1s^2\ 2s^2\ 2p^6\ 3s^2$

period 3

$_{19}K$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

$_{20}Ca$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

Both in period 4, but effective nuclear charge are differ

Mg < Ca < K

- Mg has the smallest atomic size because it has 3 shells while K and Ca have 4 shells.
  The outer electrons of Mg experience less shielding effect than K and Ca.
  Thus, the nucleus attraction towards valence electrons of Mg become stronger.

- K and Ca are in the same period. Ca has more proton number than K. So, the effective nuclear charge increases from K to Ca. The nucleus attraction towards valence electrons of Ca become stronger.
  Therefore, the atomic size of Ca is smaller than K.

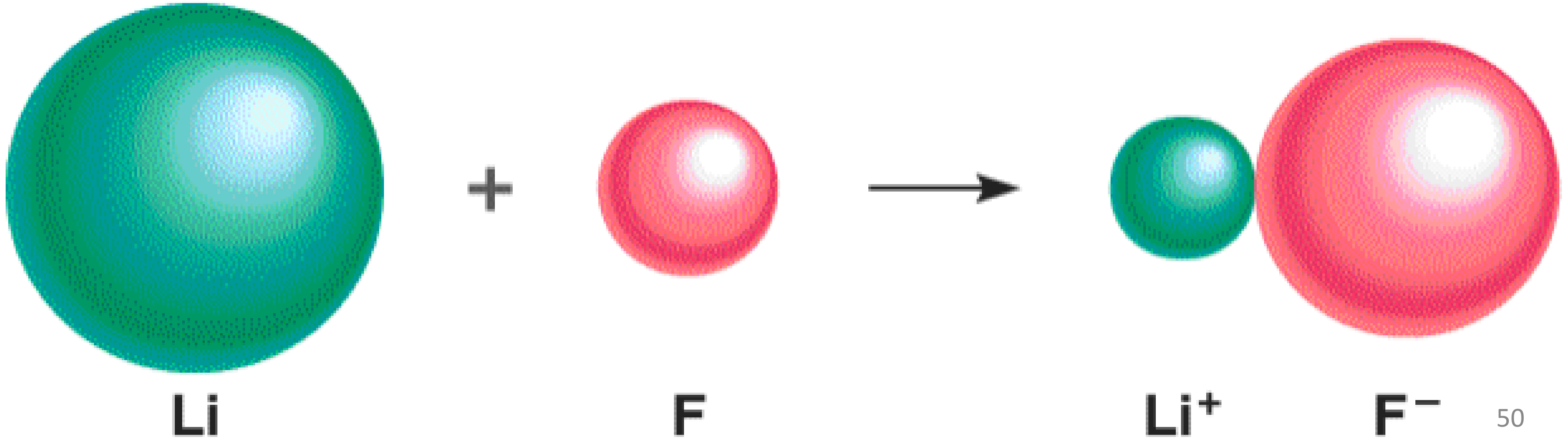
TRENDS AMONG TRANSITION ELEMENTS

A Atomic radius (pm)

- **Atomic size decreases at first but the remain fairly constant**
- **The *d* electrons fill inner orbitals, so they shield outer electrons from the increasing nuclear change much more effectively. As a result, the outer 4*s* electrons are not pulled closer**

IONIC SIZE

- Compare to their parent atoms:
  - ✔ Cations are smaller
  - ✔ Anions are larger



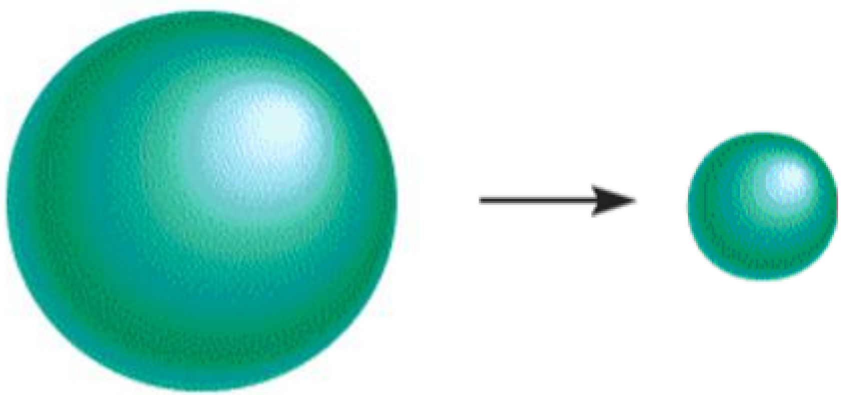
EXAMPLE:

Sodium, Na(Z = 11)

e^- configuration: $1s^2\ 2s^2\ 2p^6\ 3s^1$

Na (Z = 11) ⟶ Na^+

$1s^2\ 2s^2\ 2p^6\ 3s^1$ ⟶ $1s^2\ 2s^2\ 2p^6$



EXAMPLE:

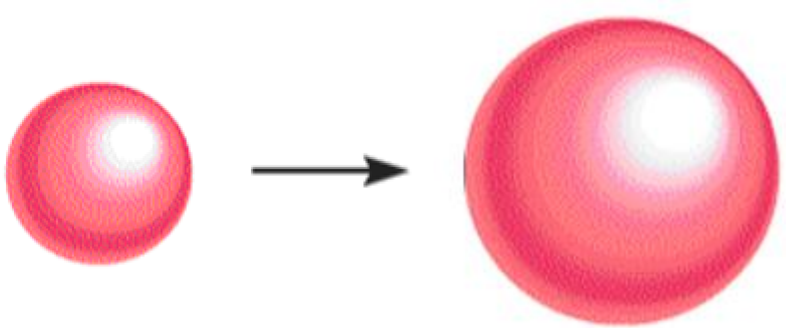
	Na	Na^+
Number of protons	11	11
Nuclear charge (*Z*)	+11	+11
Number of e^-	11	10

- When a cation forms, e^- are removed from the outer shell
- The electron cloud will shrink
- Electron–electron repulsion decreases
- The attraction of nucleus towards the remaining electron stronger
- The size of cation decreases

EXAMPLE:

F (Z = 9) ⟶ **F^-**

$1s^2\ 2s^2\ 2p^5$ **$1s^2\ 2s^2\ 2p^6$**

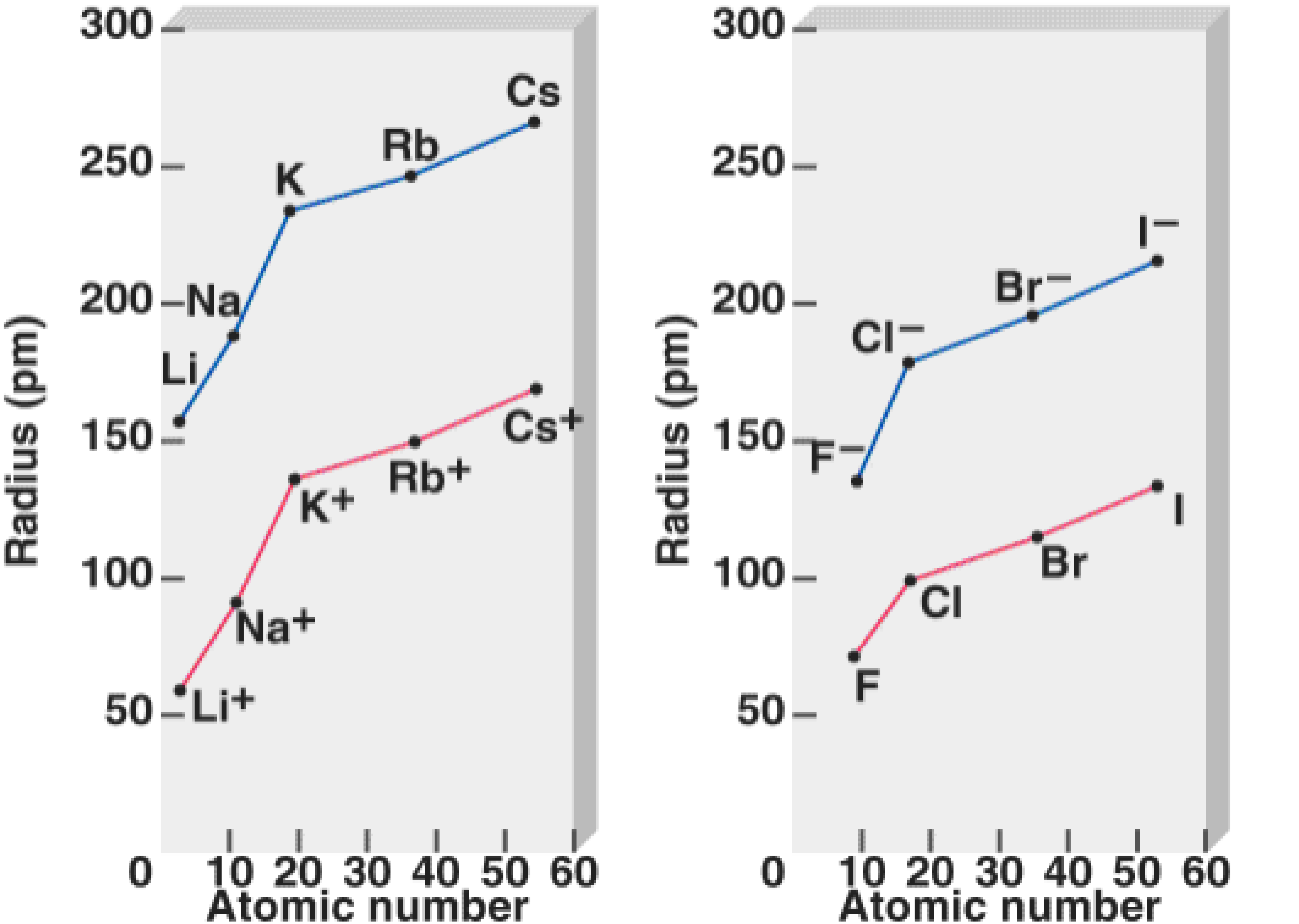


EXAMPLE:

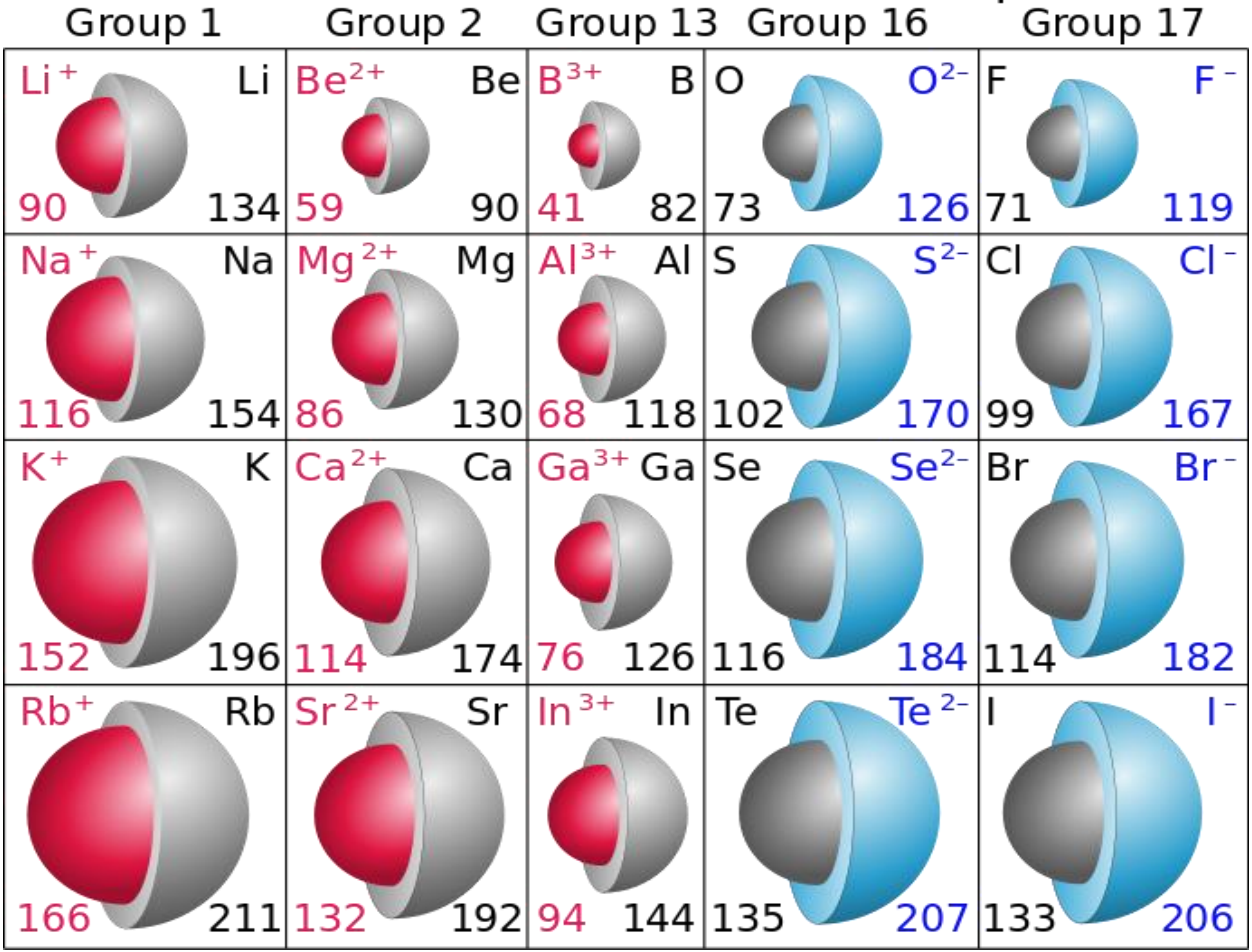
	F	F^-
Number of protons	9	9
Nuclear charge (*Z*)	+9	+9
Number of e^-	9	10

- When an anion forms, e^- are added from the outer shell
- Electron–electron repulsion increases
- Electron cloud expand
- The attraction of nucleus towards outer electron weaker
- The size of anion increases

Comparison of Atomic Radii with Ionic Radii



Sizes of atoms and their ions in pm



Example 3

Write the electronic configuration for each species. Choose the larger particle in each pair:

(a) $_{13}Al$ or Al^{3+}

(b) $_{16}S$ or S^{2-}

Answer

(a) Al or Al^{3+}

Al : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Al^{3+} : $1s^2\ 2s^2\ 2p^6$

Size: Al atom is larger than Al^{3+}.

(b) S or S^{2-}

$_{16}S$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

S^{2-} : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Size: S^{2-} atom is larger than S atom.

ISOELECTRONIC

- **Isoelectronic species are groups of atoms and ions which have the same electronic configuration.**

- **Within isoelectronic species:**

  - **The bigger the nuclear charge, the stronger the attraction of nucleus towards the remaining electrons. Thus, the smaller the species.**

EXAMPLE :

Na^{+}, Mg^{2+}, Al^{3+} and Si^{4+} are isoelectronic :

Na^{+} = $1s^{2}2s^{2}2p^{6}$

Mg^{2+} = $1s^{2}2s^{2}2p^{6}$

Al^{3+} = $1s^{2}2s^{2}2p^{6}$

Si^{4+} = $1s^{2}2s^{2}2p^{6}$

Chemistry unit

➢ Isoelectronic species with electronic configuration $1s^2\ 2s^2\ 2p^6$ (10 electron)

Species	Number of electron	Number of proton
Na^+	10	11
Mg^{2+}	10	12
Al^{3+}	10	13
Si^{4+}	10	14

- When proton number increase, effective nuclear charge increase.
- The attraction between nucleus and remaining electrons stronger.
- Therefore, the ionic radii decrease.

The ionic radii of $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$

- Isoelectronic species with electronic configuration
  - $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ (18 electron)

Species	Number of electron	Number of proton
P^{3-}	18	15
S^{2-}	18	16
Cl^{-}	18	17

- When proton number increase, effective nuclear charge increase.
- The attraction between nucleus and remaining electrons stronger.
- Therefore, the ionic radii decrease.

The ionic radii of $Cl^{-} < S^{2-} < P^{3-}$

EXAMPLE:

Size: N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}

Period 2 (N^{3-}, O^{2-}, F^{-}) **Period 3** (Na^{+}, Mg^{2+}, Al^{3+})

- **All ions have total 10 electrons (isoelectronic).**
- **When the proton number increase, the effective nuclear charge increase.**
- **Attraction between nucleus and the remaining electrons stronger.**
- **So, size of the ions decrease.**

Example 4

Give the electron configurations of these ions, and indicate which ones are isoelectronic.

(a) Ca^{2+}

(b) K^{+}

(c) O^{2-}

Given $_{20}Ca$, $_{19}K$, $_{8}O$

Answer

(a) $_{20}Ca$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

Ca^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

(b) $_{19}K$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

K^+: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

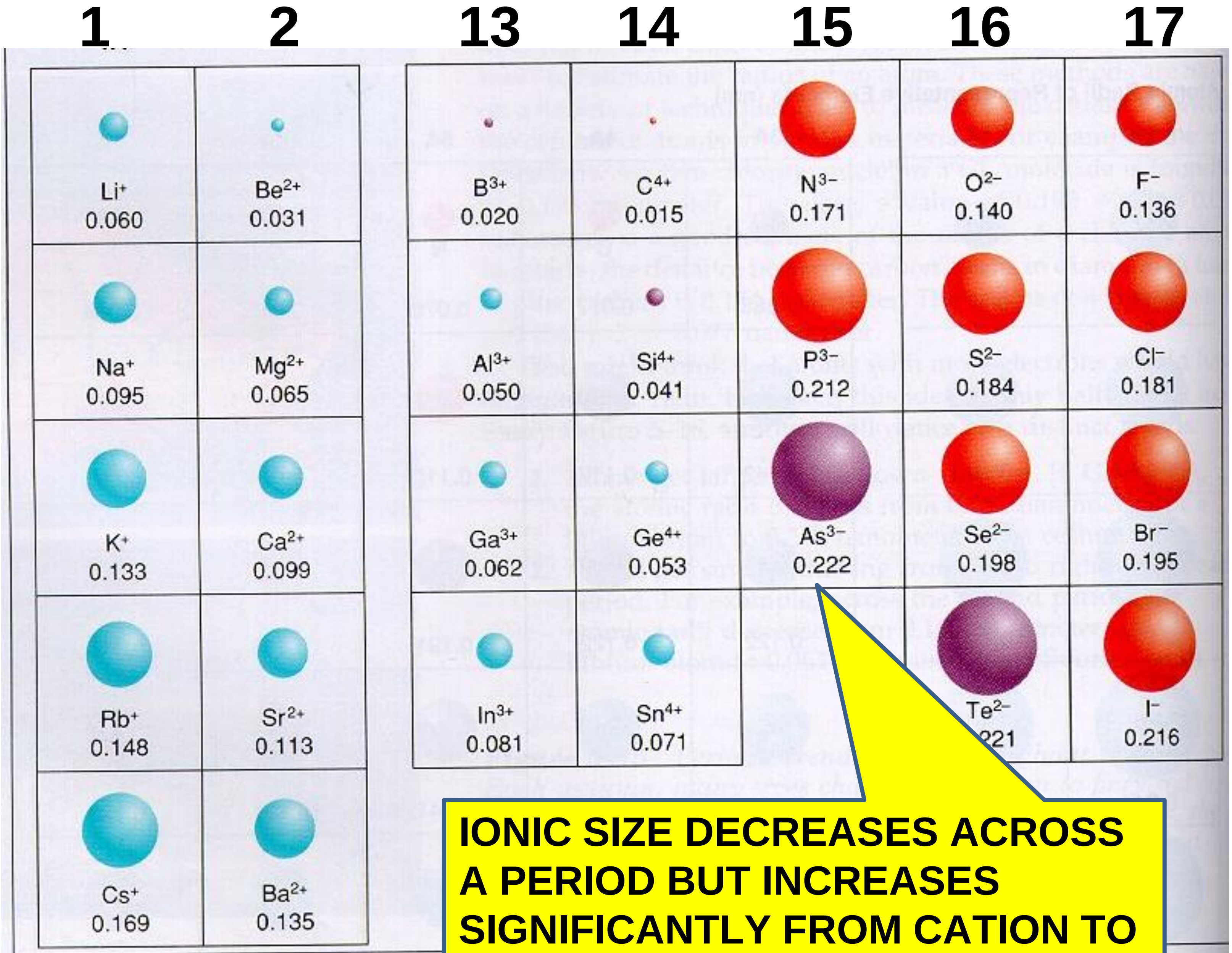
(c) $_8O$: $1s^2\ 2s^2\ 2p^4$

O^{2-}: $1s^2\ 2s^2\ 2p^6$

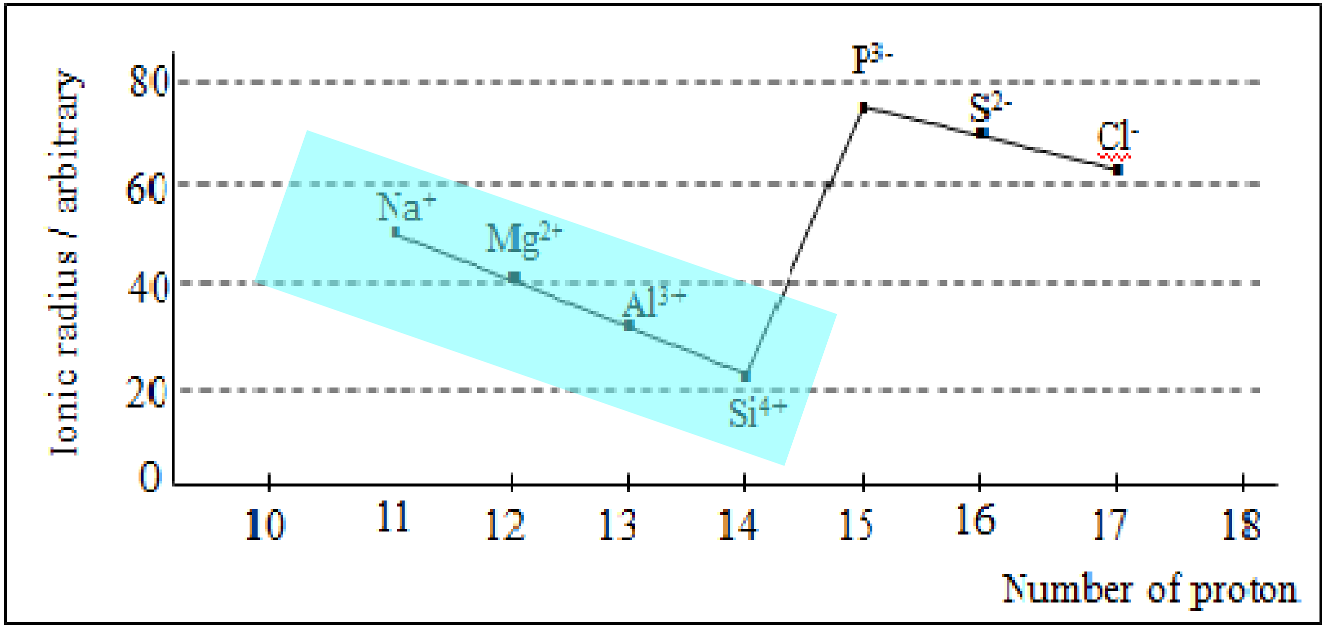
K^+ and Ca^{2+} are isoelectronic

TRENDS IN IONIC SIZE

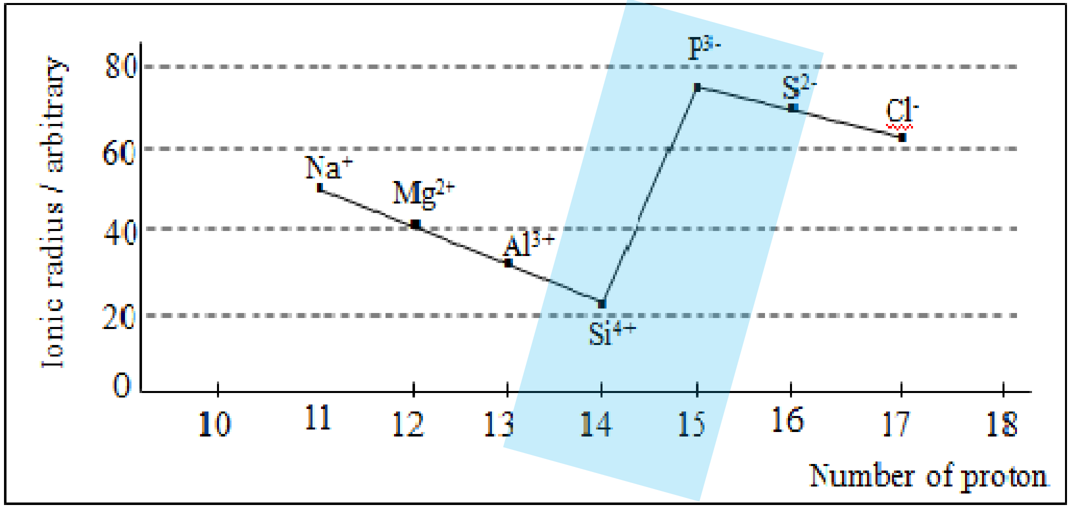
- The ionic size decreases across a period but increases significantly from cation to anion



IONIC SIZE DECREASES ACROSS A PERIOD BUT INCREASES SIGNIFICANTLY FROM CATION TO ANION



- They are **isoelectronic : Na^+ to Si^{4+} (10 electrons)** with same electronic configuration : : **$1s^2\ 2s^2\ 2p^6$**
- Moving from Na^+ to Si^{4+} , number of **protons** increase , **effective nuclear charge increase**.
- So, **nucleus–electrons attraction stronger**
- **Size of ions decreases**



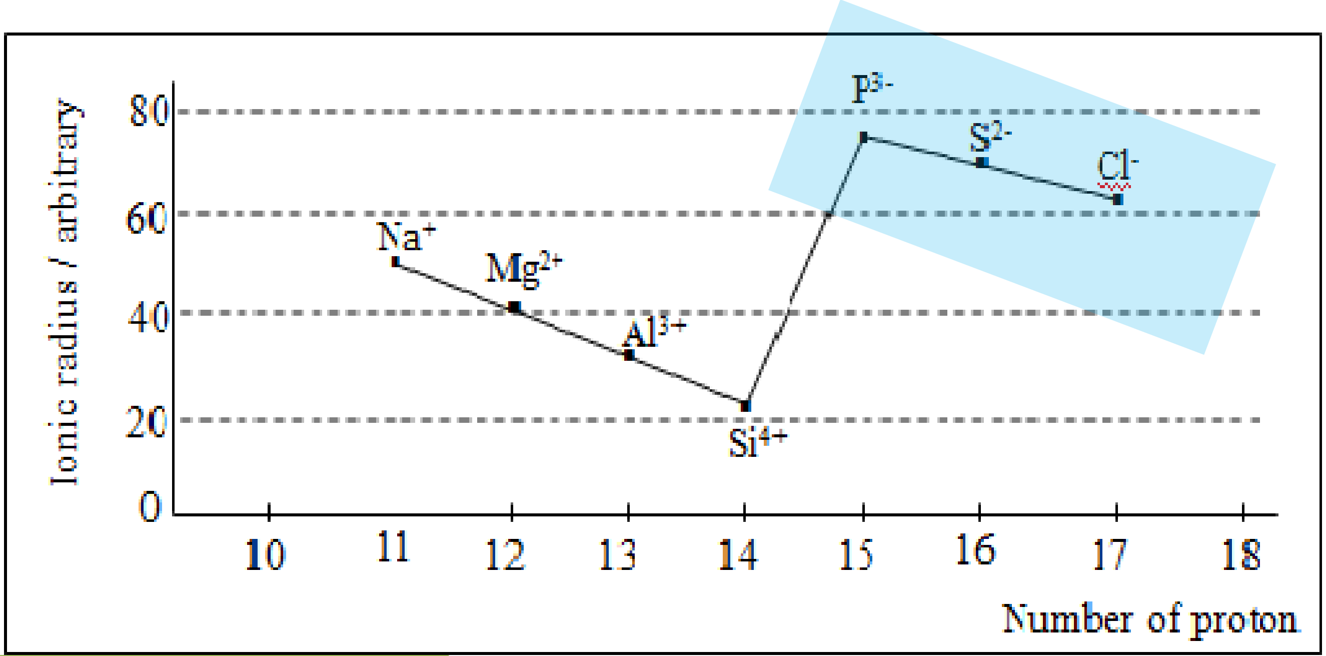
- **Great jump** in size from Si^{4+} (10e⁻) to P^{3-} (18 e⁻)

✔ Electrons are added to 3p orbital to form P^{3-} . So, the principle quantum no. (n) increases.

✔ The completely-filled inner orbital electrons have effectively shielded the outer electrons in P^{3-} from the attraction of the nucleus attraction.

✔ Electron– electron repulsion increases and nucleus–electrons attraction weaker

✔ Size increases sharply



- **From P^{3-} to Cl^{-}**

✔ **They are isoelectronic (18e) , with same electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$.**

✔ **From P^{3-} to Cl^{-}, number of protons increase, leads to increasing in effective nuclear charge .**

✔ **So, electron held more tightly to the nucleus**

✔ **Size of ions decreases**

IONISATION ENERGY

DEFINITION Ionisation energy is minimum energy required for complete removal of 1 mol of electron from 1 mol of gaseous atoms or gaseous ions.

FIRST IONISATION ENERGY (IE_1)	SECOND IONISATION ENERGY (IE_2)
Minimum energy required for complete removal of 1 mol of electron from 1 mol of gaseous atoms.	Minimum energy required for complete removal of 1 mol of electron from 1 mol of gaseous ions (charge +1).
Equation : $Atom(g) \rightarrow ion^+(g) + e^-$ ΔE=+ve	Equation : $Ion^+(g) \rightarrow ion^{2+}(g) + e^-$ ΔE =+ve
EXAMPLE: $Na(g) \rightarrow Na^+(g) + e^-$ ΔE=496 kJ/mol	EXAMPLE: $Na^+(g) \rightarrow Na^{2+}(g) + e^-$ ΔE=4560kJ/mol

Down a group; IE_1 decreases

- Down a group, the atomic size increase because the principal quantum number (*n*) and shielding effect increase

- The distance between nucleus to outer/ valence e^- increases

- The nucleus–electron attraction weaker, so easier to remove outer / valence e^-

- Less energy needed to remove e^-

❖ IE_1 decreases down the group

Across a period; IE_1 increase

- Across a period, the atomic size decreases because the proton number increases and Z_{eff} increases
- The distance of nucleus to outer / valence e^- decreases
- Nucleus–electron attraction stronger, so difficult to remove outer/valence e^-
- More energy needed to remove e^-

❖ IE_1 increases across a period

Example 5

Write the electronic configuration for each species. Rank the elements in each of the following sets in order of decreasing IE_1:

(a) $_{36}Kr$, $_{2}He$, $_{18}Ar$

(b) $_{19}K$, $_{20}Ca$, $_{37}Rb$

Answer

(a) $_{36}Kr$, $_{2}He$, $_{18}Ar$ (they are in the same group)

$_{36}Kr$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

$_{2}He$: $1s^2$

$_{18}Ar$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

In order of decreasing IE_1: **He > Ar > Kr**

Answer

(b) $_{19}K$, $_{20}Ca$, $_{37}Rb$

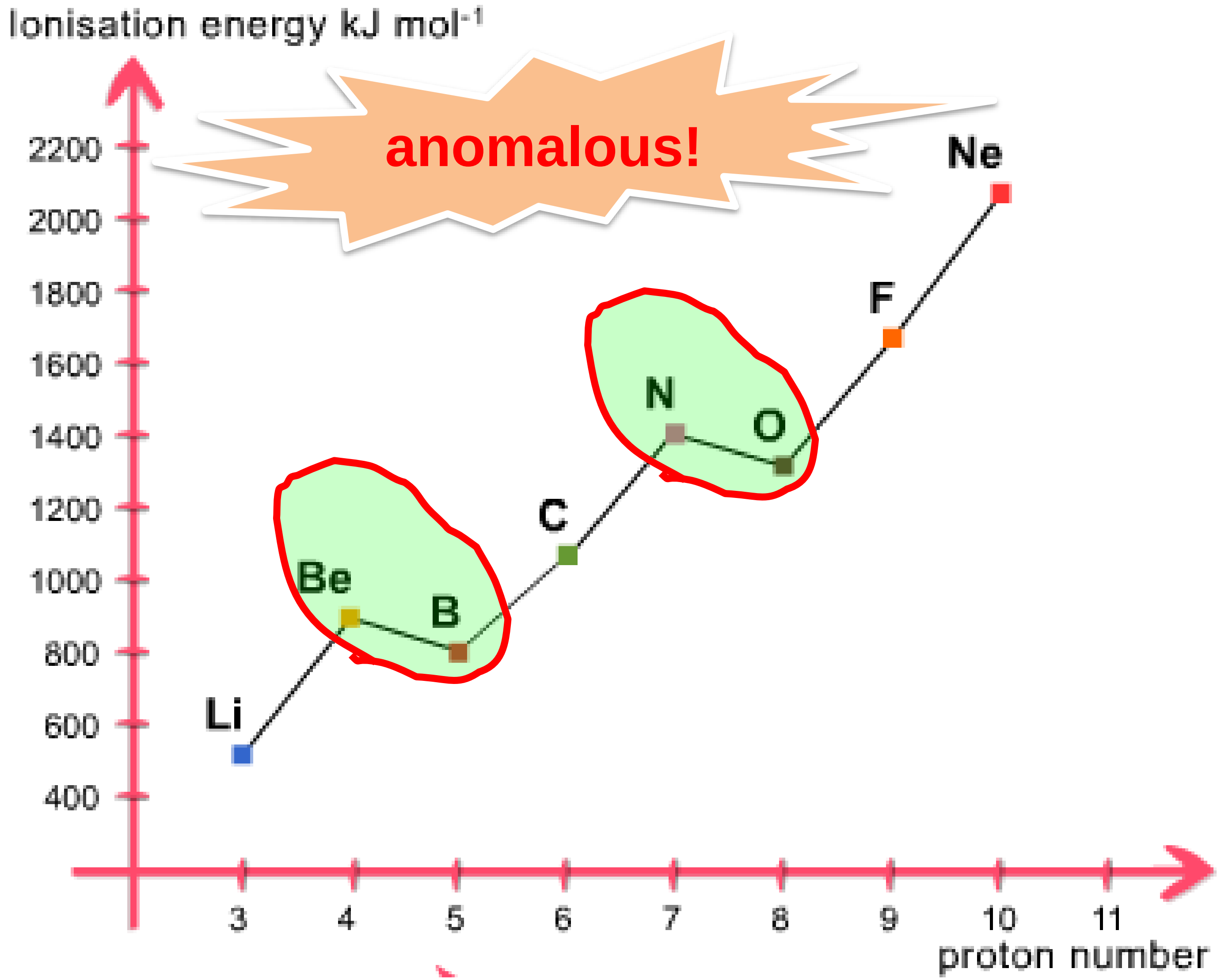
$_{19}K$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

$_{20}Ca$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

$_{37}Rb$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$

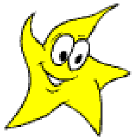
In order of decreasing IE_1: **Ca > K > Rb**

ANOMALOUS IN IONISATION ENERGY



ANOMALOUS IN IONISATION ENERGY

❑ Between Group 2 and 13:



$Be(g) \rightarrow Be^{+}(g) + 1e^{-}$

$1s^2\,2s^2 \qquad 1s^2\,2s^1$

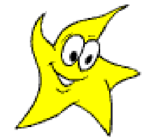
$B(g) \rightarrow B^{+}(g) + 1e^{-}$

$1s^2\,2s^2\,2p^1 \qquad 1s^2\,2s^2$

➢ IE_1 of B< Be because the the 2p subshell is at higher energy than the 2s subshell. Thus less energy is required to remove a single 2p electron in B than to remove an electron from $2s^2$ of Be.

ANOMALOUS IN IONISATION ENERGY

❑ Between Group 15 and 16:



$N(g) \rightarrow N^+(g) + 1e^-$

$1s^2\,2s^2\,2p^3$
(half-filled 2p orbital)

$O(g) \rightarrow O^+(g) + 1e^-$

$1s^2\,2s^2\,2p^4$
(partially-filled 2p orbital)

- N : electron is removed from the half filled 2p orbital, which is more stable.
- O : electron is removed the partially filled 2p orbital, which is less stable.
- More energy required to removed an electron from more stable 2p orbital of N than from less stable 2p orbital of O.

Example 6

In general, ionisation energy increase from left to right across a given period.

Aluminium and Magnesium are in same period however, Aluminium (Al), has a lower first ionization energy than magnesium (Mg).

Explain.

Answer

$$Mg(g) \rightarrow Mg^{+}(g) + 1e^{-}$$

$$1s^2\ 2s^2\ 2p^6\ 3s^2 \qquad 1s^2\ 2s^2\ 2p^6\ 3s^1$$

$$Al(g) \rightarrow Al^{+}(g) + 1e^{-}$$

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1 \qquad 1s^2\ 2s^2\ 2p^6\ 3s^2$$

- Mg loses an electron from the completely filled 3s which is more stable than an electron of the partially filled 3p orbital in Al. Removing an electron from the completely filled orbital needs more energy.
- The electron in 3p orbital of Al is higher in energy, hence less stable and is easier to be removed
- Therefore, the first ionization energy Al is less than Mg.

IONISATION ENERGY

1. We can deduce the **electronic configuration** of an element and **its position** in the periodic table based on **successive ionisation energy.**
2. **IE** always **increase** in the following order:

$$IE_1 < IE_2 < IE_3 < IE_4 < IE_5 < IE_6 \ldots\ldots\ldots\ldots$$

EXAMPLE: : Be (Z = 4) $1s^2\ 2s^2$

IE_1 = 900 kJ/mol

IE_2 = 1760 kJ/mol

IE_3 = 14850 kJ/mol

IE_4 = 21910 kJ/mol

SUCCESSIVE IE

$IE_1 < IE_2 < IE_3 < IE_4 < IE_5 < IE_6$ ……….

- Successive ionisation energies (IE_1, IE_2, and so on) of a a given element increase because each electron is removed away from an ion with a progressively higher positive charge

- A careful examination of the successive ionisation energy value will lead us to notify the sudden increase (drastic increase).

SUCCESSIVE IE

EXAMPLE:

First e^- removed from 2s

$Be(g) \rightarrow Be^+(g) + e^-$ $IE_1 = 900$ kJ/mol

$1s^2\ 2s^2$ $1s^2\ 2s^1$

Second e^- removed from 2s

$Be^+(g) \rightarrow Be^{2+}(g) + e^-$ $IE_2 = 1\ 760$ kJ/mol

$1s^2\ 2s^1$ $1s^2$

There is sudden increase of IE

Third e^- removed from 1s (inner shell)

$Be^{2+}(g) \rightarrow Be^{3+}(g) + e^-$ $IE_3 = 14\ 850$ kJ/mol

$1s^2$ $1s^1$

Fourth e^- removed from 1s

$Be^{3+}(g) \rightarrow Be^{4+}(g) + e^-$ $IE_4 = 21\ 910$ kJ/mol

$1s^1$

	Electronic configuration	First	Second	Third	Fourth	Sixth	First
Li	$1s^2\ 2s^1$	520	7297	11810	-	-	-
Be	$1s^2\ 2s^2$	900	1760	14850	21910	-	-
B	$1s^2\ 2s^2\ 2p^1$	800	2430	3659	25020	32810	-

For Li : sudden increase is between IE_1 to IE_2 because the second electron is removed from 1s orbital, inner shell.

For Be : sudden increase is between IE_2 to IE_3 because the third electron is removed from 1s orbital, inner shell.

For B : sudden increase is between IE_3 to IE_4 because the fourth electron is removed from 1s orbital, inner shell.

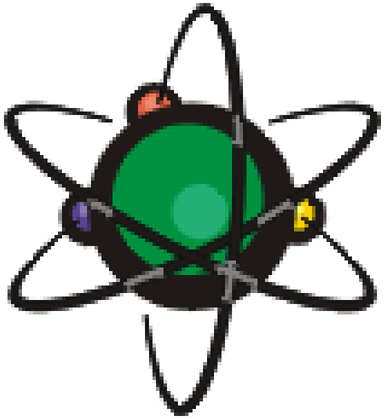
- Analysing the data of successive ionisation energy, will enable us to figure out the position of element in the Periodic Table by looking for the sudden increase.

- A very large increase in IE indicates the removal of a particular electron involves an electron from an inner shell, which has a stable noble gas electron configuration

IDENTIFYING AN ELEMENT FROM SUCCESSIVE IE

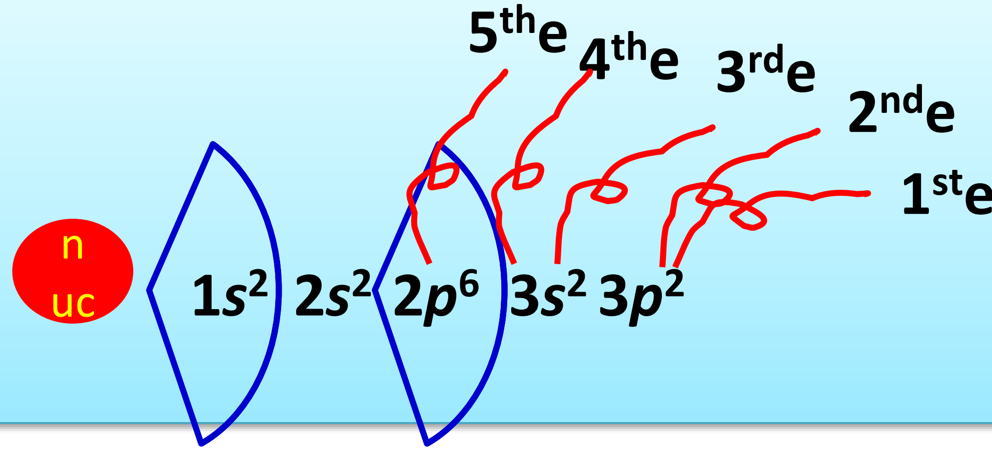
From the data, we can deduce:

- ✔ Number of **valence e$^-$**
- ✔ **Group** number of the element



EXAMPLE:

Si (Z = 14)



786 kJ/mol	IE_1: 1st e^- removed from 3*p* subshell
1580 kJ/mol	IE_2: 2nd e^- removed from 3*p* subshell
3230 kJ/mol	IE_3: 3rd e^- removed from 3*s* subshell
4360 kJ/mol	IE_4: 4th e^- removed from 3*s* subshell
16100 kJ/mol	IE_5: 5th e^- removed from 2*p* subshell ,inner shell

Sudden increase

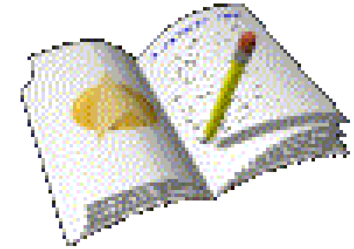
EXAMPLE:

Si (Z = 14) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

- **Sudden increase in ionisation energy from IE_4 to IE_5**
- **The 5th electron is removed from the inner shell which has a stable noble gas electron configuration and closer to nucleus**
- **Much greater energy needed to remove the e^-**

- **Valence e^- = 4 Group = 14**
- **Valence e^- configuration = $3s^2 3p^2$**

Example 7



Deduce the group number of the element with the following ionisation energies (in kJ/mol) and write its valence electron configuration.

IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
1012	1903	2910	4956	6278	22,230

Answer

IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
1012	1903	2910	4956	6278	22,230

- **Sudden increase in ionisation energy from IE_5 to IE_6.**

- **The 6th electron is removed from the inner shell which has a stable noble gas electron configuration and closer to nucleus.**

- **Much greater energy needed to remove the electron.**

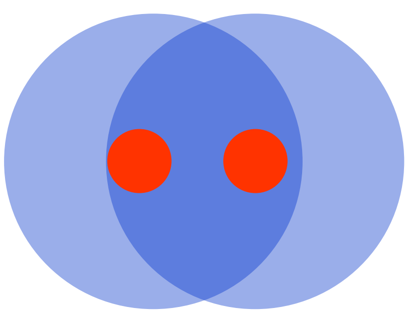
- **Valence electrons = 5 Group = 15**

- **Valence electrons configuration = ns^2 np^3**

ELECTRONEGATIVITY (EN)

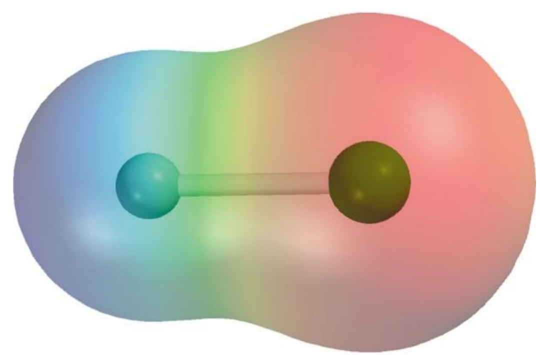
DEFINITION **Electronegativity is a relative ability of a bonded atom to attract the shared electrons toward itself in a chemical bond.**

H:H



H_2

H:F



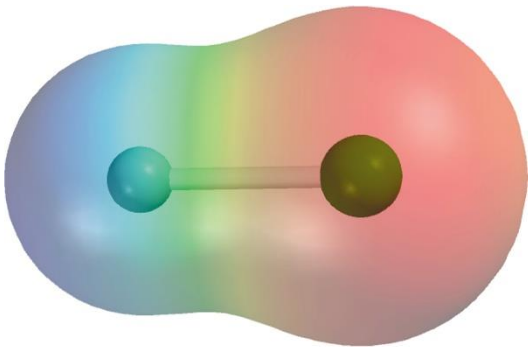
HF

EXAMPLE:

$\delta+$ $\delta-$

H–F

H : F



HF

- The e^- spend more time closer to F

- F is more electronegative than H

- Cause F ends of the bond partially negative and H end partially positive

- H–F ☞ polar bond

TRENDS IN ELECTRONEGATIVITY

- Across a period, nuclear charge ↑, atomic size ↓, so ability of an atom to attract the shared electrons toward itself increase , thus electronegativity ↑

- Down a group, no. of shell (*n*)↑ or shielding effect ↑, atomic size↑, so ability of an atom to attract the shared electrons toward itself decrease, thus electronegativity ↓

- Nonmetals are more electronegative than metal

- Highest EN value (4.0) assigned to F

EXAMPLE:

Note: ( ) ☞ electronegativity scale

- **Period 2:**

Li	Be	B	C	N	O	F
(1.0)	(1.5)	(2.0)	(2.5)	(3.0)	(3.5)	(4.0)

→ **electronegativity increase**

- **Period 3:**

Na	Mg	Al	Si	P	S	Cl
(0.9)	(1.2)	(1.5)	(1.8)	(2.1)	(2.5)	(3.0)

→ **electronegativity increase**

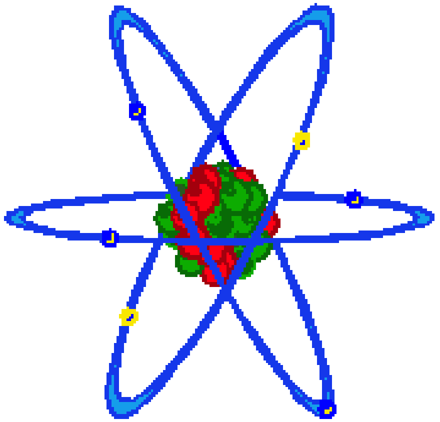
EXAMPLE:

Note: ( ) ☞ electronegativity scale

- Group 17:

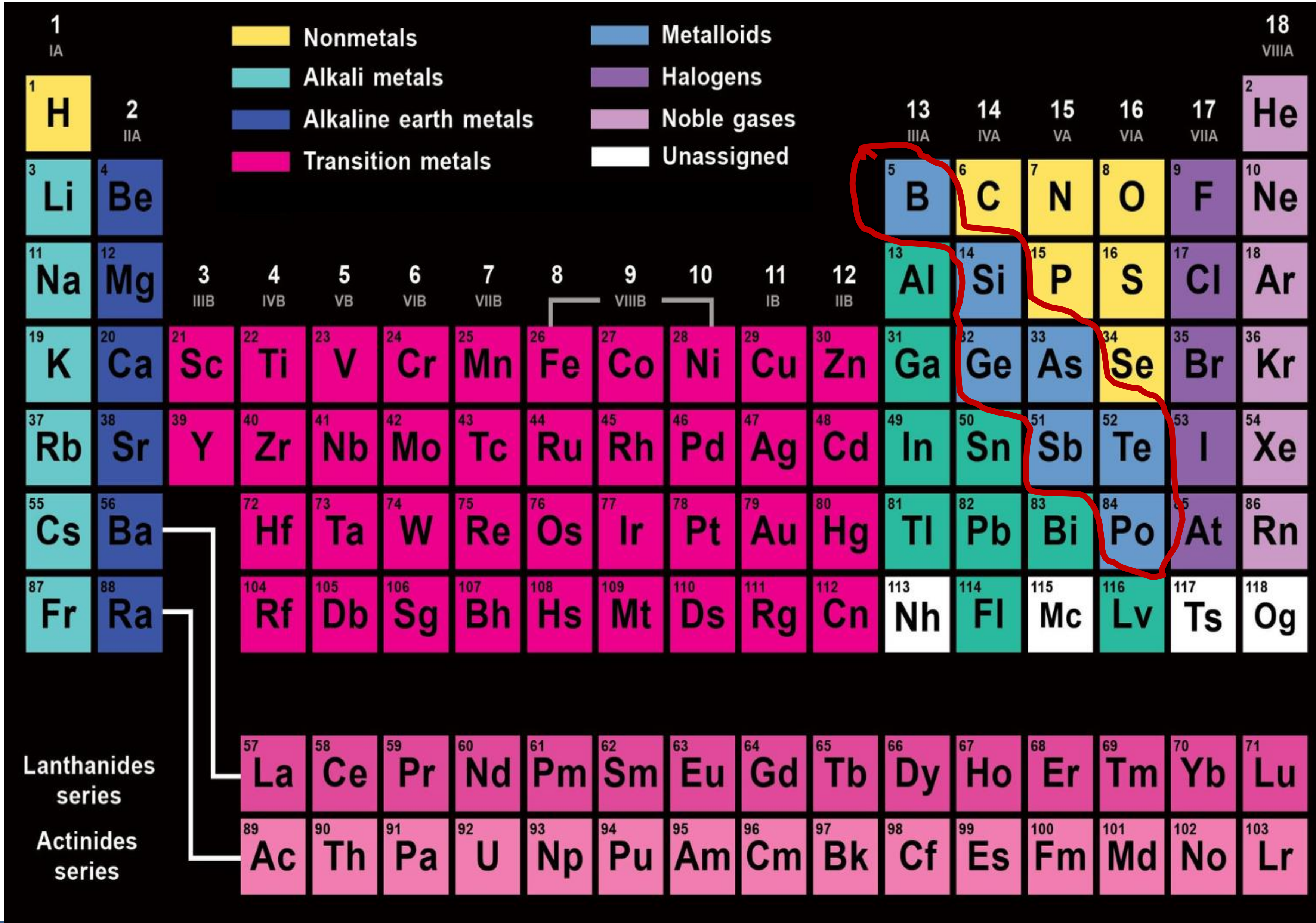
F (4.0)
Cl (3.0)
Br (2.8)
I (2.5)

Down a group electronegativity decrease



Increasing Metallic Character

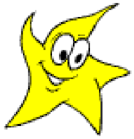
Increasing Metallic Character



ACID–BASE PROPERTIES OF OXIDES ACROSS PERIOD 3

Element	Na	Mg	Al	Si	P	S	Cl
Formula of oxides	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_4O_{10}**	**SO_3**	**Cl_2O_7**
Types of compounds	**Ionic Compounds**			**Giant Network structure**	**Molecular covalent compounds**		
Chemical properties	**Basic oxides**		**Amphoteric**	**Acidic**			
Interparticles forces	**During melting and boiling process, the substances need to overcome the strong ionic bond**			**The substance need to overcome the infinite number of strong covalent bonds**	**To boil or melt, the substance need to overcome the intermolecular forces (weak van der Waals forces)**		

Main–group metal oxides



- When react with oxygen, Na and Mg form **basic oxide**

  - These oxides are **ionic compounds**

  - Na_2O react with water to form base NaOH

$$Na_2O(s) \quad + \quad H_2O(l) \quad \rightarrow \quad 2NaOH(aq)$$

  - MgO is insoluble in water, however, it does react with acids to produce salt and water

$$MgO(s) \quad + \quad 2HCl(aq) \rightarrow \quad MgCl_2(aq) \quad + \quad H_2O(l)$$

Main–group metal oxides

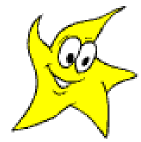
- When react with oxygen, Al form **amphotheric (act as base and acid) oxide**
- Al_2O_3 is an ionic compounds
- Al_2O_3 act as acid :

$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$$

- Al_2O_3 act as base :

$$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$$

Metalloid oxides



- When burn in oxygen, Si form acidic oxide, SiO_2

- SiO_2 has a giant network structure which contain infinite number of strong covalent bonds.

- SiO_2 is insoluble in water, however, it does react with bases to produce salt and water

$$SiO_2\,(s) + NaOH\,(aq) \rightarrow Na_2SiO_3(aq) + H_2O\,(l)$$

Nonmetal oxides

- When burn in oxygen, P, S and Cl form acidic oxide
  - These oxides are covalent compounds
  - In water, they acts as acids, producing H^+ and reacts with bases

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$$

$$Cl_2O_7(l) + H_2O(l) \rightarrow 2HClO_4(aq)$$

Thanks!

For Attention

See You The Next Chapter

End Slide

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