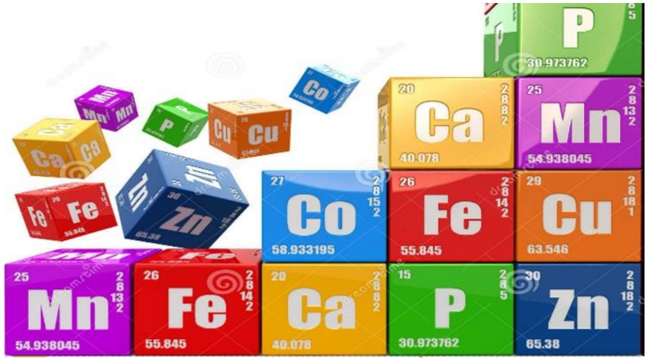
CHAPTER 3

PERIODIC TABLE

Periodic Table

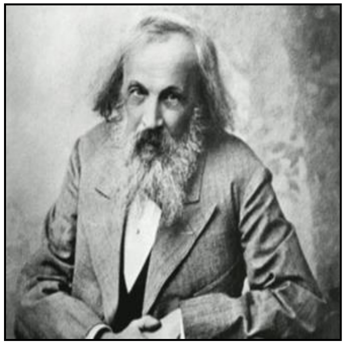
- **3.1: Classification of Elements**
- **3.2: Periodicity**



Who invented Periodic Table?

The scientists involved are : Antoine Lavoisier, Johann W. Dobereiner, John Newlands, Lothar Meyer, Dimitri Mendeleev and H.J. G. Moseley.

Lothar Meyer - properties of element were in complied a periodic table of 56 elements based on regular repeating pattern of physical properties.



Dimitri Mendeleev - published the first version of the periodic table in 1869. Arranged elements according to increasing atomic masses.

Henry Moseley - Rearrange the elements in the periodic table by their atomic numbers.

THE MODERN PERIODIC TABLE

Consists of

4 BLOCKS

s,p,d & f

7 rows

Called

PERIOD

18 vertical columns

Called

GROUP

107 elements

Period, Group & Block (*s,p,d,f*)

- The element in periodic table are arranged in **order of increasing proton number.**

- The position of elements can be determined by using **electronic configuration.**

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	57-71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Period, Group & Block (*s*,*p*,*d*,*f*)

- A vertical column of elements is called a **group**.
- A horizontal row of elements is known as a **period**.

PERIOD

GROUP

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	57-71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo
		57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
		89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Group

- **Elements in the same group have the same number of valence electrons, thus same chemical properties.**

Group 1 -2

group no = number of valence electron

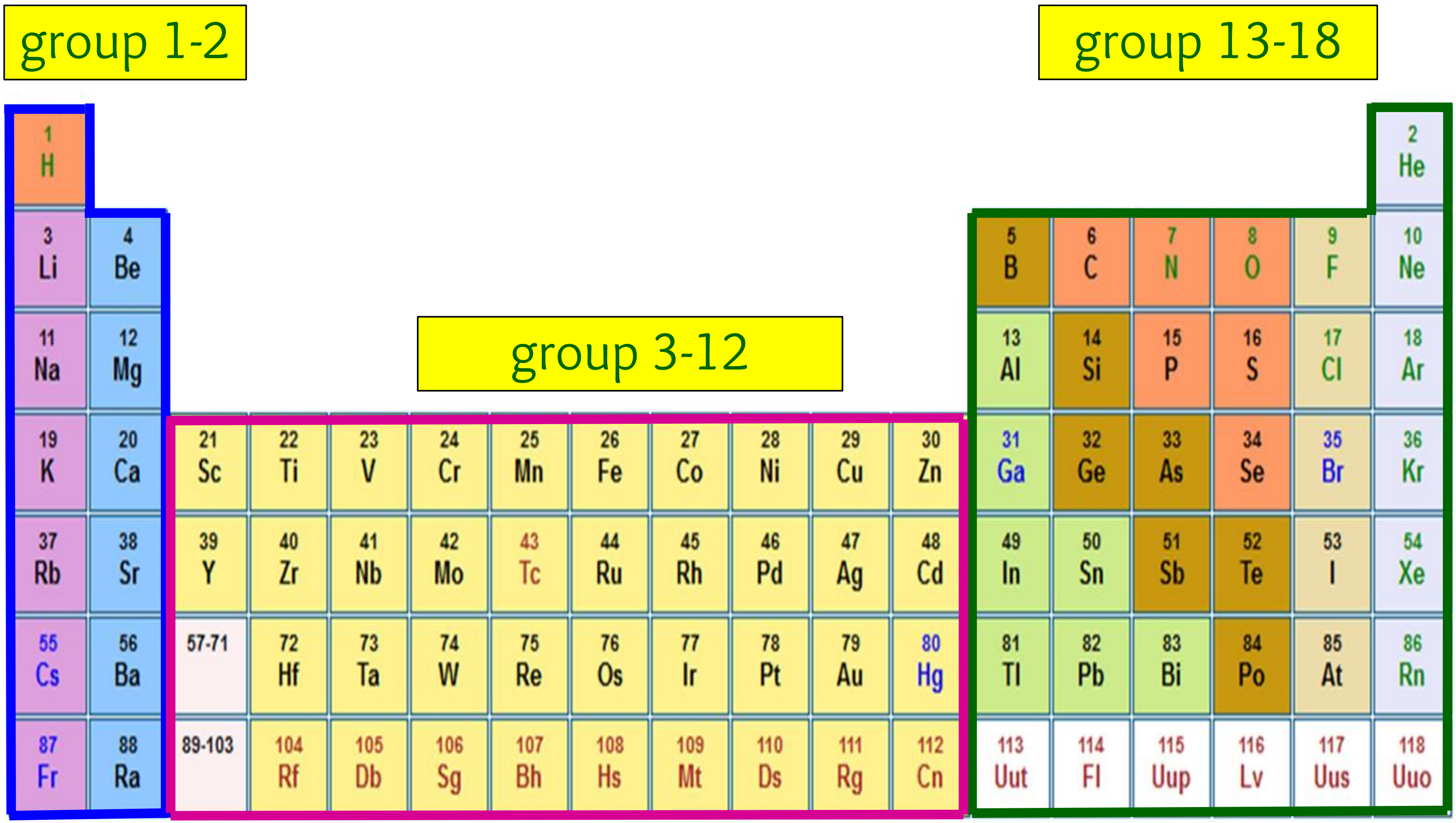
Group 13-18

group no = number of valence electron + 10

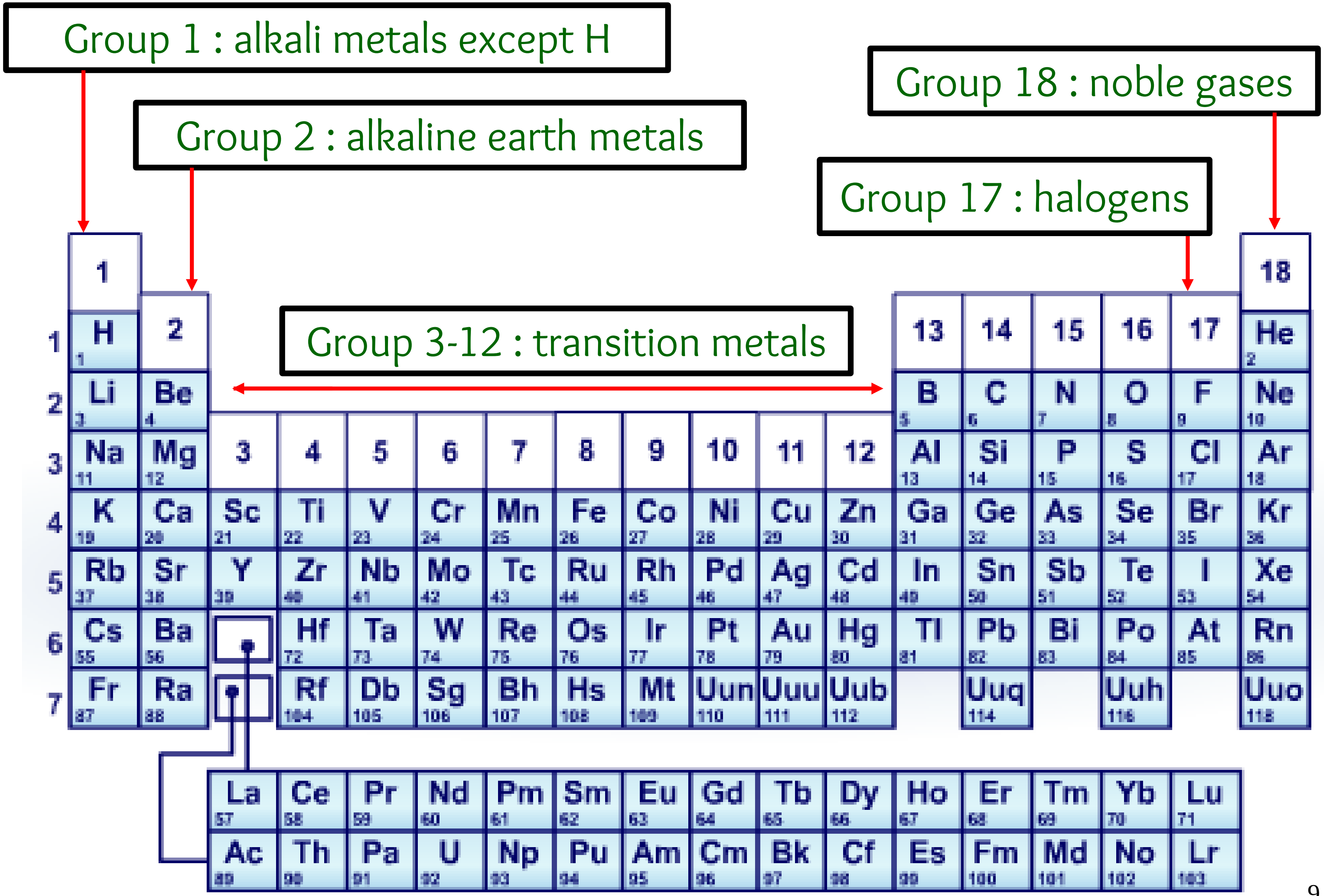
Group 3-12

group no = number of valence electron of $ns^2 (n-1)d^1$ to $ns^2 (n-1)d^{10}$

Group


group 1-2
group 13-18
group 3-12
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
57-71
72 Hf
73 Ta
74 W
75 Re
76 Os
77 Ir
78 Pt
79 Au
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
87 Fr
88 Ra
89-103
104 Rf
105 Db
106 Sg
107 Bh
108 Hs
109 Mt
110 Ds
111 Rg
112 Cn
113 Uut
114 Fl
115 Uup
116 Lv
117 Uus
118 Uuo

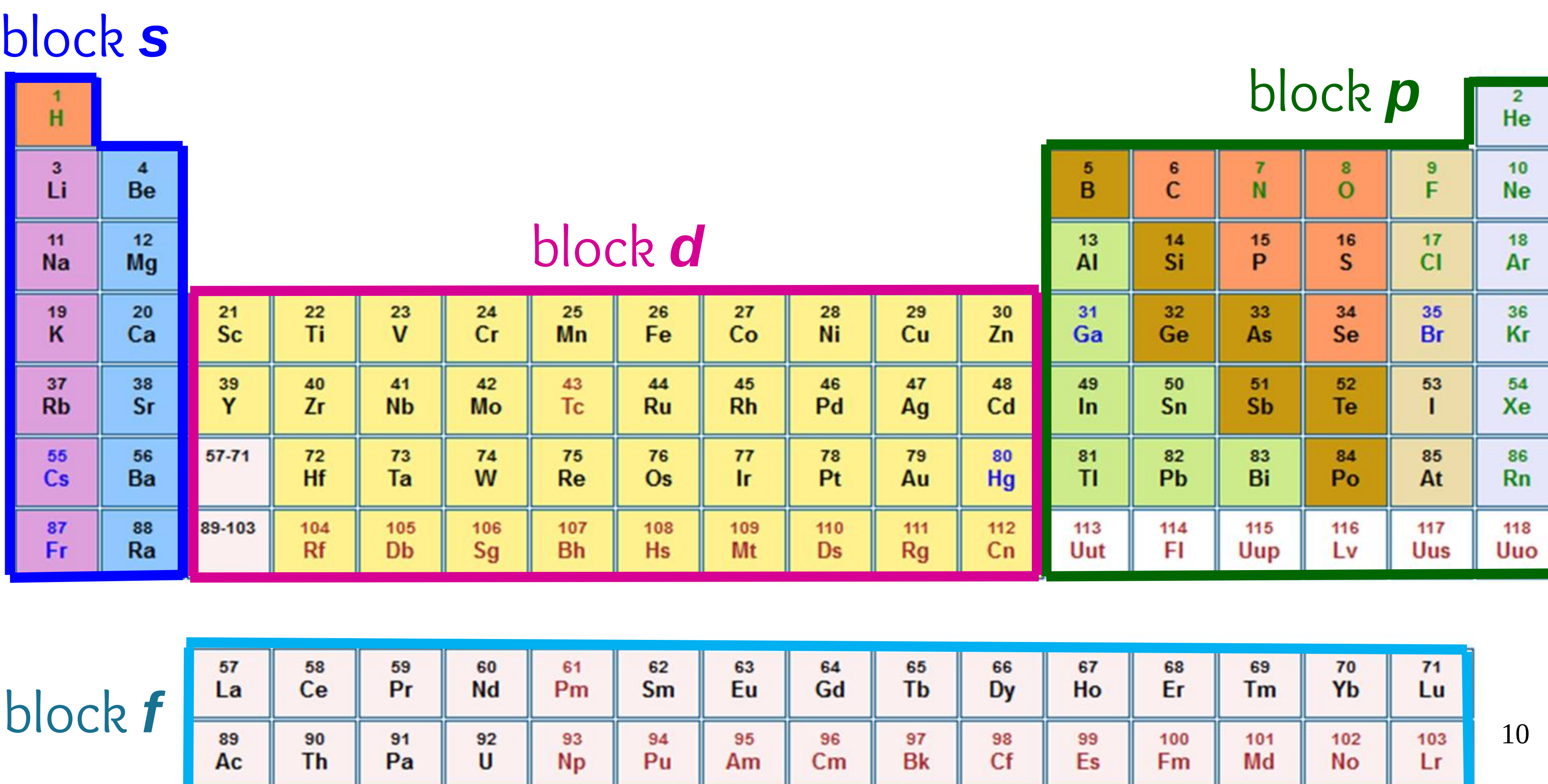
Group



Group 1 : alkali metals except H
Group 2 : alkaline earth metals
Group 18 : noble gases
Group 17 : halogens
Group 3-12 : transition metals
1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17 18
1 H 1 He 2
2 Li 3 Be 4 B 5 C 6 N 7 O 8 F 9 Ne 10
3 Na 11 Mg 12 Al 13 Si 14 P 15 S 16 Cl 17 Ar 18
4 K 19 Ca 20 Sc 21 Ti 22 V 23 Cr 24 Mn 25 Fe 26 Co 27 Ni 28 Cu 29 Zn 30 Ga 31 Ge 32 As 33 Se 34 Br 35 Kr 36
5 Rb 37 Sr 38 Y 39 Zr 40 Nb 41 Mo 42 Tc 43 Ru 44 Rh 45 Pd 46 Ag 47 Cd 48 In 49 Sn 50 Sb 51 Te 52 I 53 Xe 54
6 Cs 55 Ba 56 Hf 72 Ta 73 W 74 Re 75 Os 76 Ir 77 Pt 78 Au 79 Hg 80 Tl 81 Pb 82 Bi 83 Po 84 At 85 Rn 86
7 Fr 87 Ra 88 Rf 104 Db 105 Sg 106 Bh 107 Hs 108 Mt 109 Uun 110 Uuu 111 Uub 112 Uuq 114 Uuh 116 Uuo 118
La 57 Ce 58 Pr 59 Nd 60 Pm 61 Sm 62 Eu 63 Gd 64 Tb 65 Dy 66 Ho 67 Er 68 Tm 69 Yb 70 Lu 71
Ac 89 Th 90 Pa 91 U 92 Np 93 Pu 94 Am 95 Cm 96 Bk 97 Cf 98 Es 99 Fm 100 Md 101 No 102 Lr 103

Block

- Elements in the periodic table can be classified into four blocks according to their valence electronic configuration.
- These blocks are block ***s & p*** (main block), ***d*** and ***f***.



Block

- The periodic table is divided into blocks according to their valence electronic configuration.

	Block		
	s	***p***	***d***
Group	1 & 2	13 –18	3 – 12
Valence electronic configuration	ns^1 to ns^2	$ns^2\ np^1$ to $ns^2\ np^6$	$ns^2\ (n\text{-}1)d^1$ to $ns^2\ (n\text{-}1)d^{10}$

Group no = number of valence electron

Example 1 : The elements in Group 1

elements	electronic configuration	valence electronic configuration	valence electrons
$_{3}Li$	$1s^2\ 2s^1$	$2s^1$	1
$_{11}Na$	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$3s^1$	1
$_{19}K$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	$4s^1$	1

Note : valence electrons are electron in the outer shell

Group no = number of valence electron

Example 2 : The elements in Group 2

elements	electronic configuration	valence electronic configuration	valence electrons
$_{4}Be$	$1s^2\ 2s^2$	$2s^2$	2
$_{12}Mg$	$1s^2\ 2s^2\ 2p^6\ 3s^2$	$3s^2$	2
$_{20}Ca$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	$4s^2$	2

Group no = number of valence electron + 10

Example 3 : The elements in Group 13

elements	electronic configuration	valence electronic configuration	valence electrons
$_{5}B$	$1s^2\ 2s^2\ 2p^1$	$2s^2\ 2p^1$	3
$_{13}Al$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	$3s^2\ 3p^1$	3
$_{31}Ga$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$	$4s^2\ 4p^1$	3

Group no = number of valence electron + 10

Example 4 : The elements in Group 17

element	electronic configuration	valence electronic configuration	valence electrons
$_{9}F$	$1s^2\ 2s^2\ 2p^5$	$2s^2\ 2p^5$	7
$_{17}Cl$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	$3s^2\ 3p^5$	7
$_{35}Br$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$	$4s^2\ 4p^5$	7

Group no = number of valence electron of $ns^2 (n-1)d^1$ to $ns^2 (n-1)d^{10}$

Example 5 :

elements	electronic configuration	valence electronic configuration	valence electrons	group
$_{21}Sc$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$	$4s^2 3d^1$	3	3
$_{23}V$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$	$4s^2 3d^3$	5	5
$_{30}Zn$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$	$4s^2 3d^{10}$	12	12

Group no = number of valence electron of $ns^2 (n\text{-}1)d^1$ to $ns^2 (n\text{-}1)d^{10}$

Example 6 :

$_{39}Y$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^1$

valence electrons = 3

valence electronic configuration = $5s^2\ 4d^1$

group = 3

Period

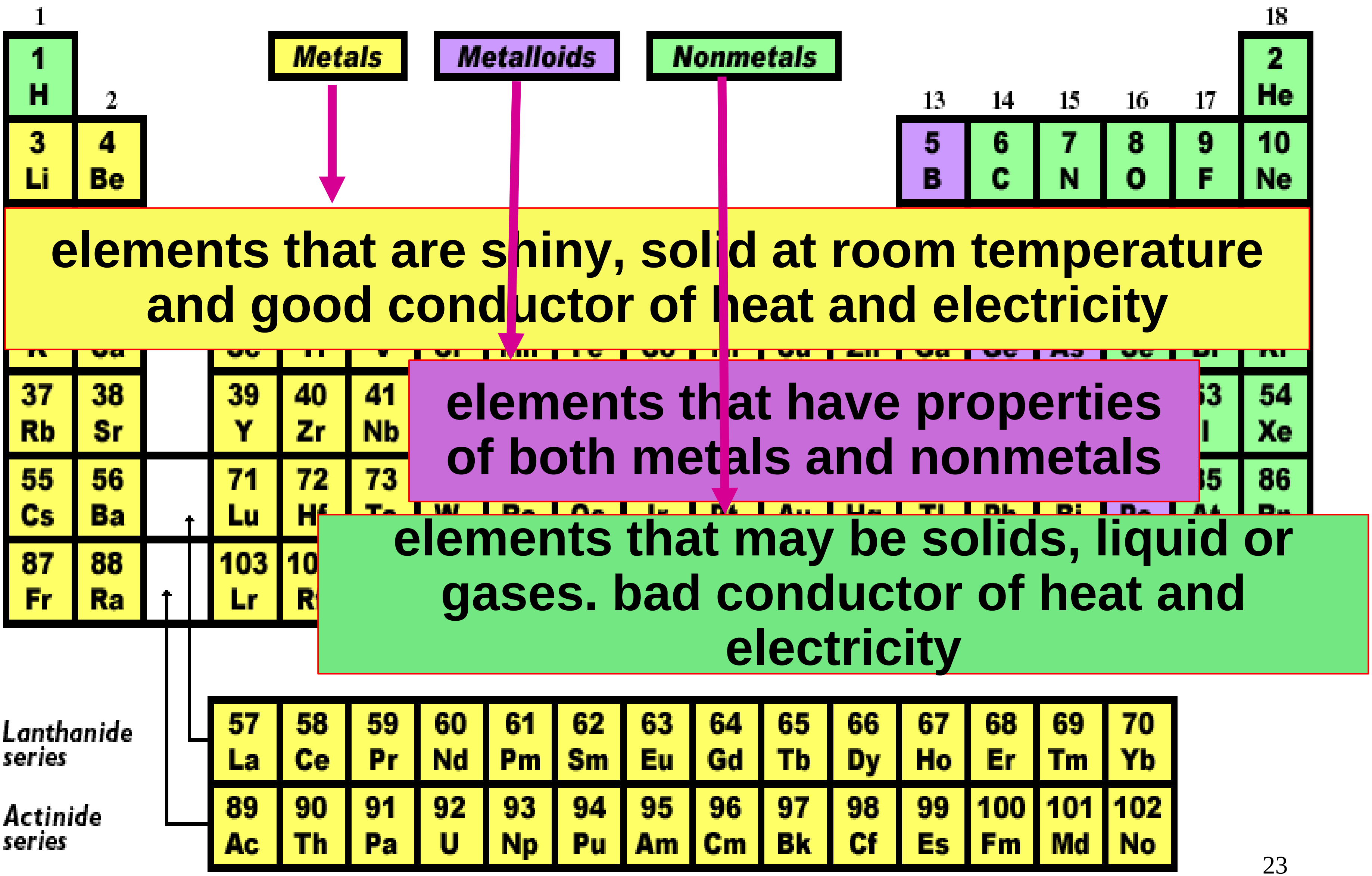
- All of the elements in a period have the same number of **highest principle quantum number, *n*.**
- Are numbered from 1 to 7.

Example :

elements	electronic configuration	higher ***n***	period
$_{3}Li$	$1s^2\ 2s^1$	2	2
$_{18}Ar$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	3	3
$_{19}K$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	4	4
$_{21}Sc$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$	4	4

**** elements of the same period have the same number of electron shells***

The position of metals, metalloids & non-metals



elements that are shiny, solid at room temperature and good conductor of heat and electricity

elements that have properties of both metals and nonmetals

elements that may be solids, liquid or gases. bad conductor of heat and electricity

CHAPTER 3.2 : OVERVIEW

PERIODICITY

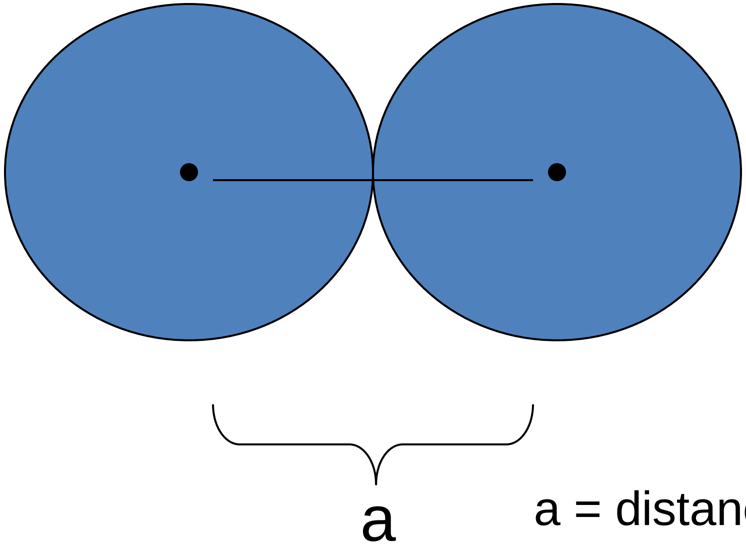
- Variation in atomic radii
- Isoelectronic species
- Variation ionic radii across Period 2 and 3
- Ionisation Energy
- Successive Ionisation Energy
- Electronegativity
- Acid base Character of Oxides

PERIODICITY

Periodicity is the periodic trend in properties of elements.

Atomic radii

Radius, r = half of the distance between the nuclei of two adjacent identical atoms.



$\boxed{\textbf{Radius, r = a/2 (Å)}}$

a = distance between two nuclei adjacent atoms

Variation in Atomic Radii

The atomic radius of an element is determined by two factors

Effective nuclear charge, Z_{eff}

Shielding/Screening effect

Effective Nuclear Charge, Z_{eff}

is the net positive charge experienced by valence electrons.

$$\boxed{Z_{eff} = Z - \boldsymbol{\sigma}}$$

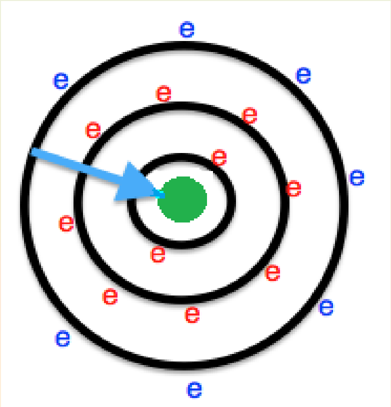
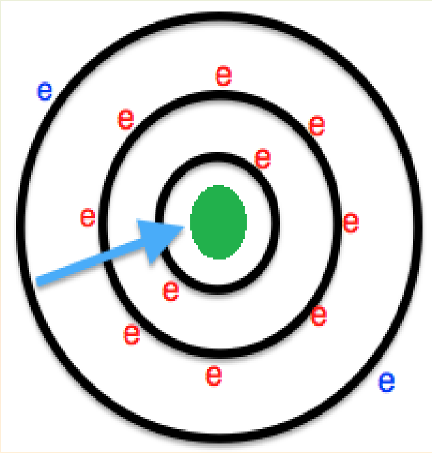
Z = number of proton

σ = number of electrons filled at the inner shell

- ✓ When effective nuclear charge, Z_{eff} increases, the nucleus attraction towards electrons become stronger.
- ✓ The nucleus pulls the outer electrons closer.
- ✓ The atomic radius becomes smaller.

Effective Nuclear Charge, Z_{eff}

Example :

$_{17}Cl$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ 	$_{12}Mg$: $1s^2\ 2s^2\ 2p^6\ 3s^2$ 
Z_{eff} : 17-10 = +7 The atomic radius of Cl is smaller because the nucleus pulls the outer electrons closer with a charge of +7	Z_{eff} : 12-10 = +2 The atomic radius of Mg is bigger than Cl because the nucleus can only pull the outer electrons closer with a charge of +2

e : outer electrons   e : inner electrons

Effective Nuclear Charge, Z_{eff}

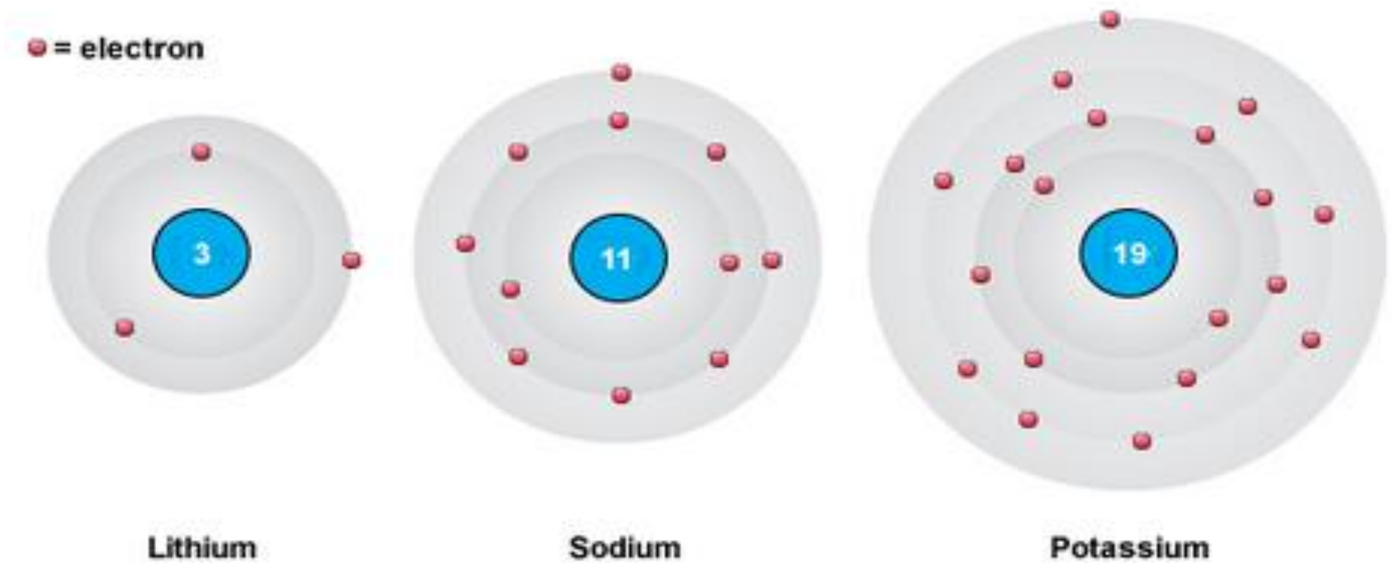
How to calculate the effective nuclear charge, Z_{eff}?

Example:

Elements	F (Fluorine)	Al (Aluminium)	K (Potassium)
Electronic configuration	$1s^2\ 2s^2\ 2p^5$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Proton number, Z	9	13	19
Number of electrons at the inner orbital, **σ**	2	10	18
Z_{eff}	+7	+3	+1

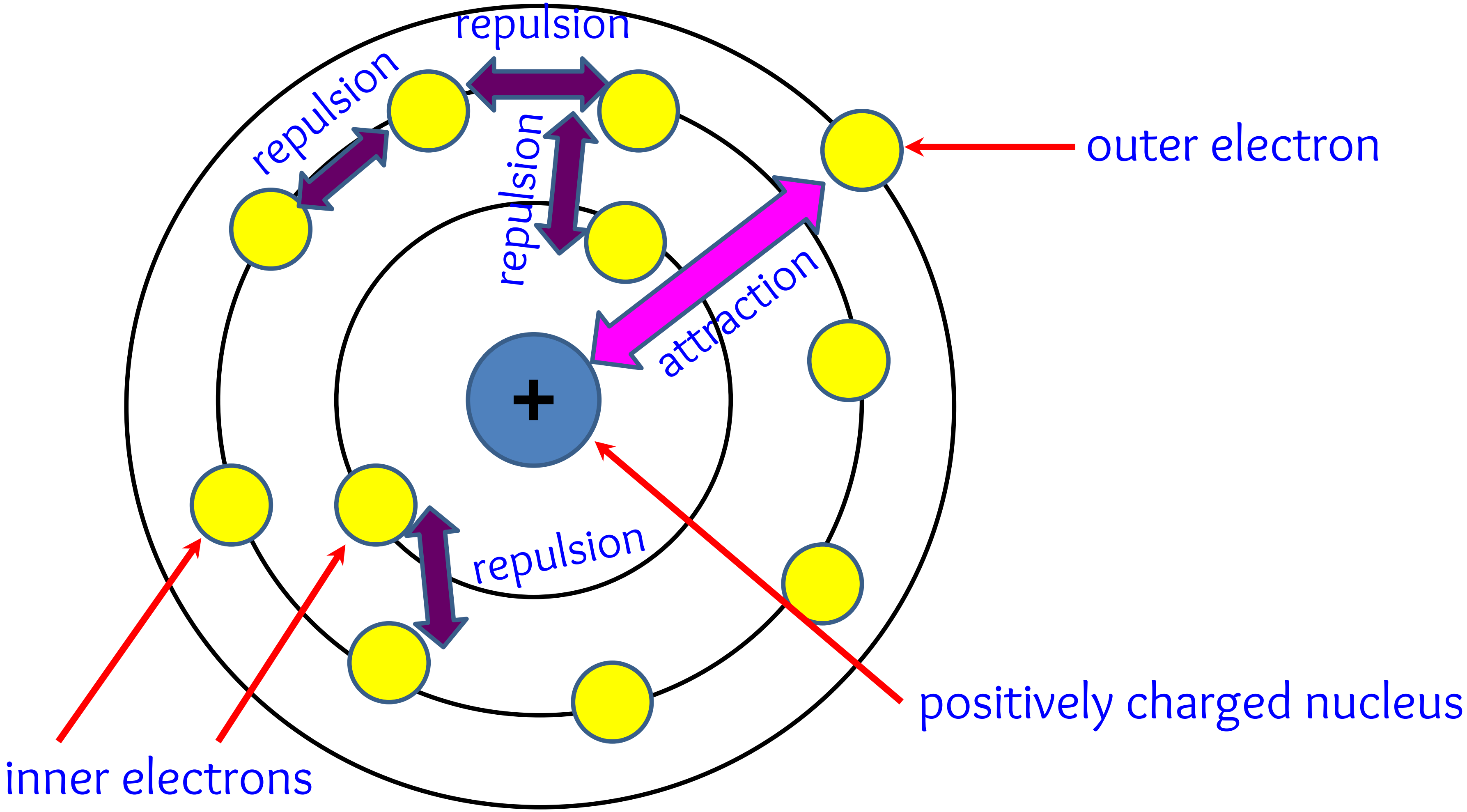
Shielding Effect

- ✓ Also known as the screening effect.
- ✓ Is caused by the mutual repulsion between electrons of inner shell and the electrons occupying valence shell.
- ✓ It also occurs between electrons in the same shell but is less effective compared to that of electrons in the different shells.



ANSWER :	EXPLANATION
(a) K > Ca > Ga $_{19}K$ = [Ar] $4s^1$ $_{20}Ca$ = [Ar] $4s^2$ $_{31}Ga$ = [Ar] $3d^{10}$ $4s^2$ $4p^1$	Elements are of the same period. Across the period from K to Ga, proton no. increase, ∴ Z_{eff} increase, hence atomic size is reduced.
(b) Rb > Br > Kr $_{35}Br$ = [Ar] $3d^{10}$ $4s^2$ $4p^5$ $_{36}Kr$ = [Ar] $3d^{10}$ $4s^2$ $4p^6$ $_{37}Rb$ = [Kr] $5s^1$	***Br & Kr are element of the same period. Br has less proton no. than Kr, ∴ Br has lower Z_{eff} compared to Kr. Hence, larger in size.*** ***Rb has a higher energy level (period 5). Inner shell increase, ∴ shielding effect increase. Atomic size bigger.***

Shielding Effect



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

Down group 1,

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

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

K $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

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

- ✓ No. of inner shells increases
- ✓ shielding effect increase
- ✓ Weaker attractive force between proton in the nucleus and the valence ê
- ✓ Valence ê are loosely held
- ✓ Size of atoms increase

ANALOGY:

Atomic Size is like the Spots on a Giraffe's Neck

The **increasing atomic size as you go down a group** in the periodic table is **similar to the pattern of increasing size** **shown by the** **spots on a giraffe's neck as you go from its head down to the body**.

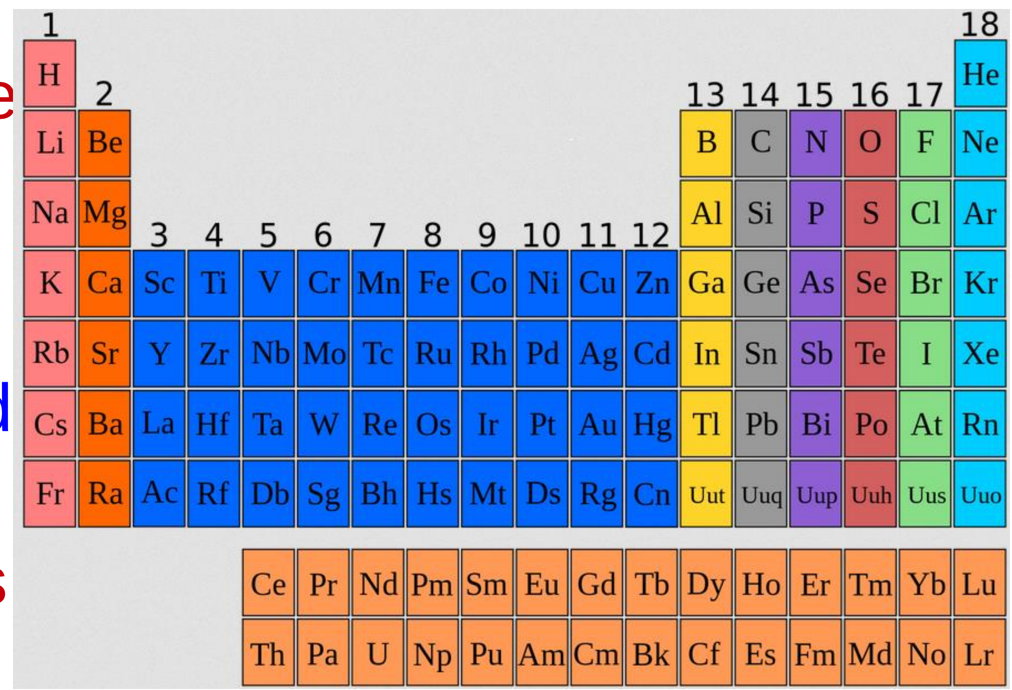
Variation in Atomic Radii

Across a period :

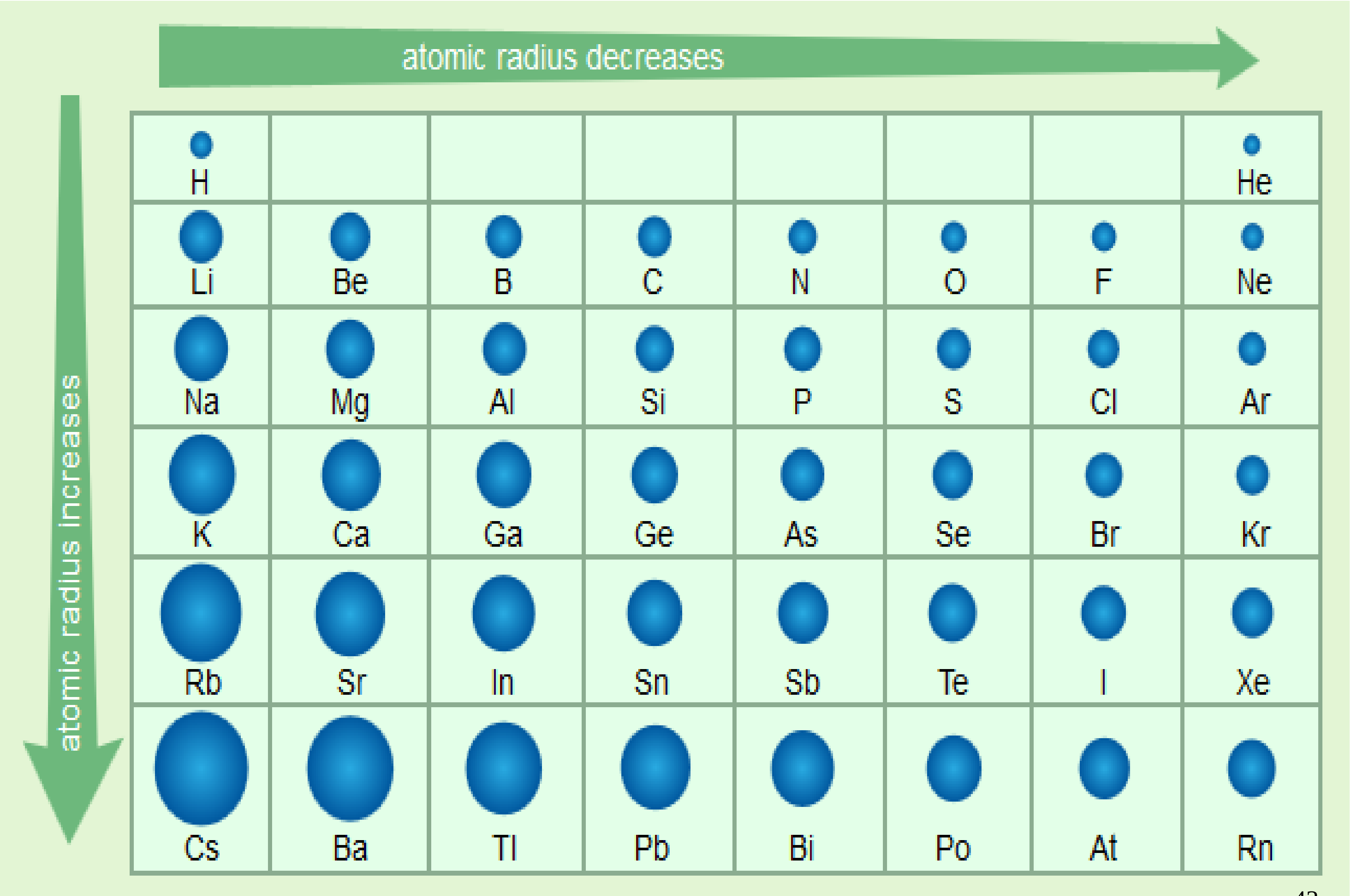
- proton number increases
- effective nuclear charge, Z_{eff} increases
- attraction between valence electrons and nucleus stronger
- atomic size smaller

Down a group :

- no of shell, *n* increase
- shielding effect increases
- attraction between valence electrons and nucleus weaker.
- atomic size increases



Variation in Atomic Radii



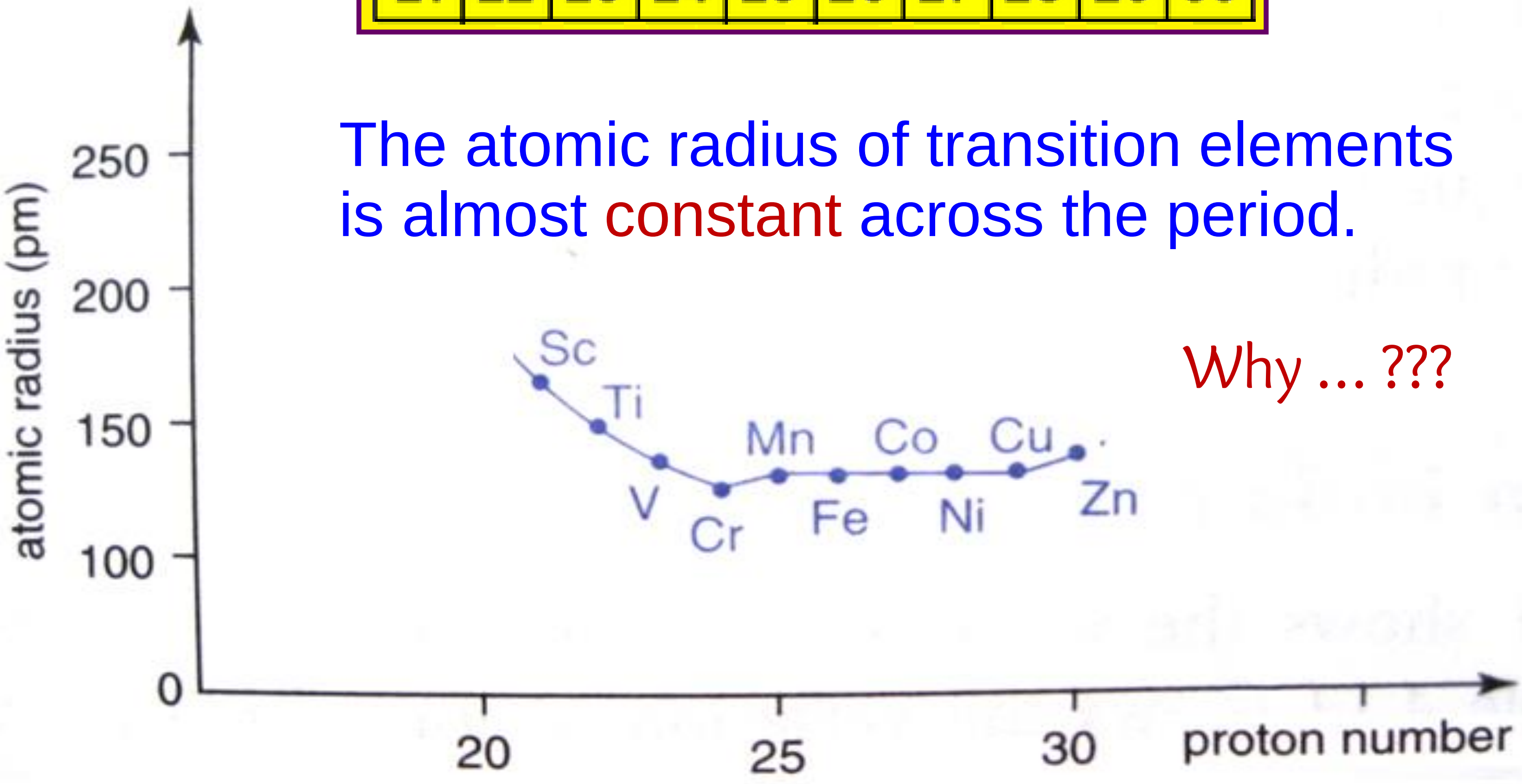
Variation in Atomic Radii

Across The First Row of Transition Elements

Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
21	22	23	24	25	26	27	28	29	30

The atomic radius of transition elements is almost constant across the period.

Why … ???



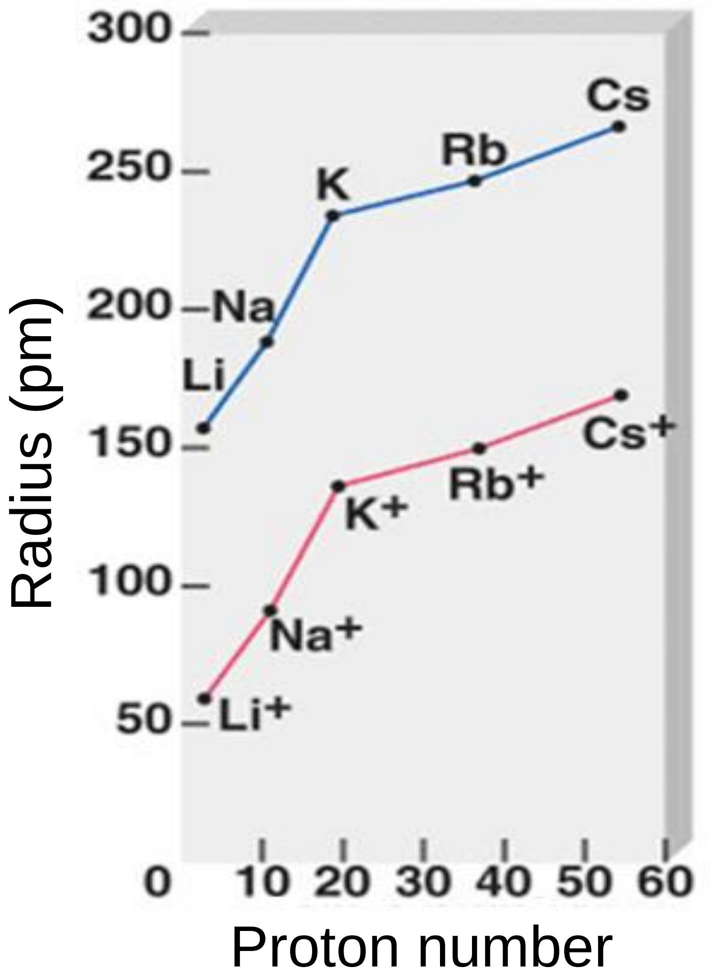
Variation in Atomic Radii

Atomic Radius/Radii of Transition Elements

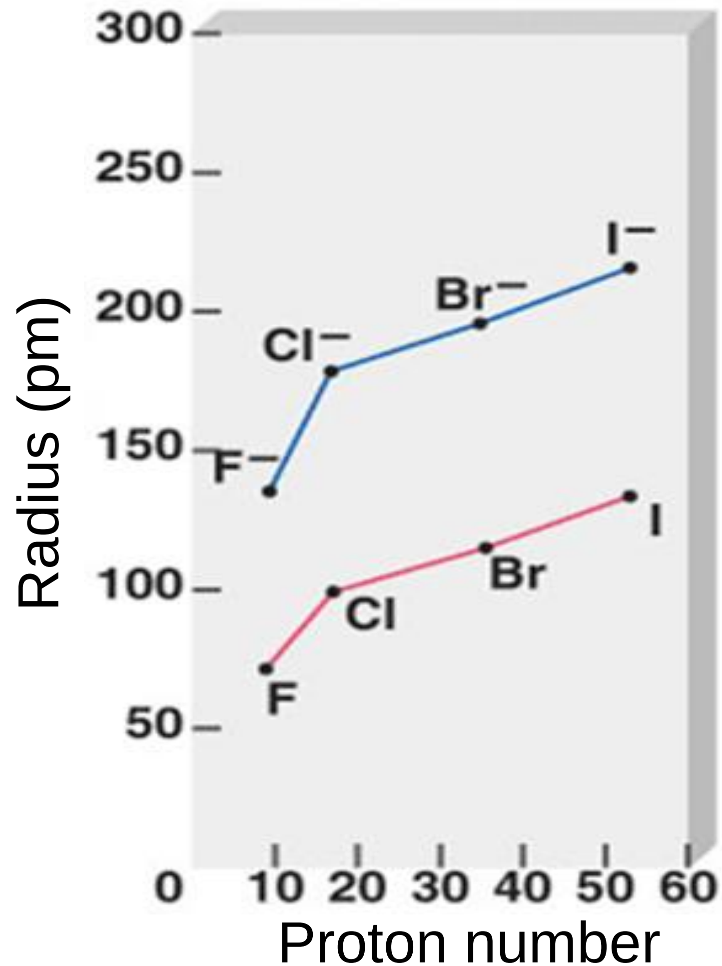
- The atomic sizes do not change significantly across the period.
- This is because the electrons are added to the inner d-orbitals.
- The 3d electrons shield the outer 4s electrons from the attractive force of the nucleus
- Thus, the increase of nuclear charge (no. of proton) is cancelled by the increase of shielding effect (caused by additional inner electrons)
- Therefore, the size remains relatively constant.

Comparison Atomic Radii & Its Ionic Radii

Observe the graphs ...



atomic radius > cation
WHY??



atomic radius < anion
WHY??

Comparison Atomic Radii & Its Cation

Example…

Neutral atom

Number of protons : 11

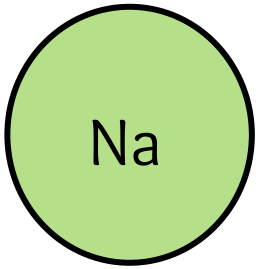
Number of electrons : 11

Cation (+ve ion)

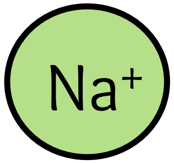
Number of proton : 11

Number of electron : 10

Comparison Atomic Radii & Its Cation



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

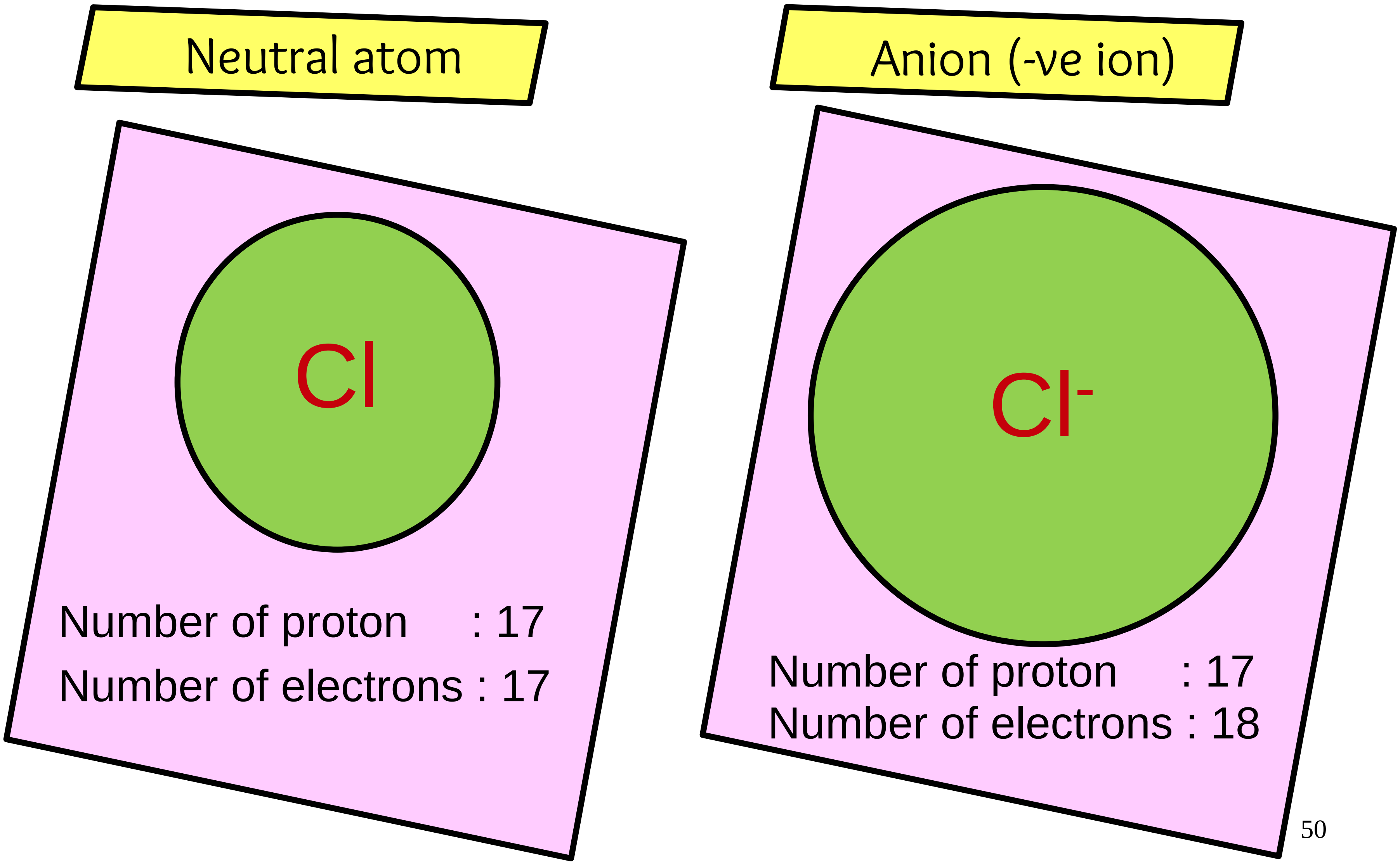


$_{11}Na^+ : 1s^2\ 2s^2\ 2p^6$

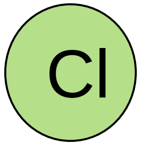
- Cation is formed when an atom loses its valence electron.
- Neutral atom (Na) & cation (Na^+) have same no. of proton, but Na^+ ion contains less electrons/no of shell than its neutral atom.
- Therefore, for Na^+ ion the attractive forces between nucleus and remaining electrons is greater than in the neutral atom.
- Thus, the size of cation is smaller than its neutral atom.

Comparison Atomic Radii & Its Anion

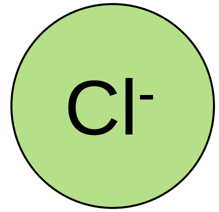
Example…



Comparison Atomic Radii & Its Anion



$_{17}Cl$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$



$_{17}Cl^-$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

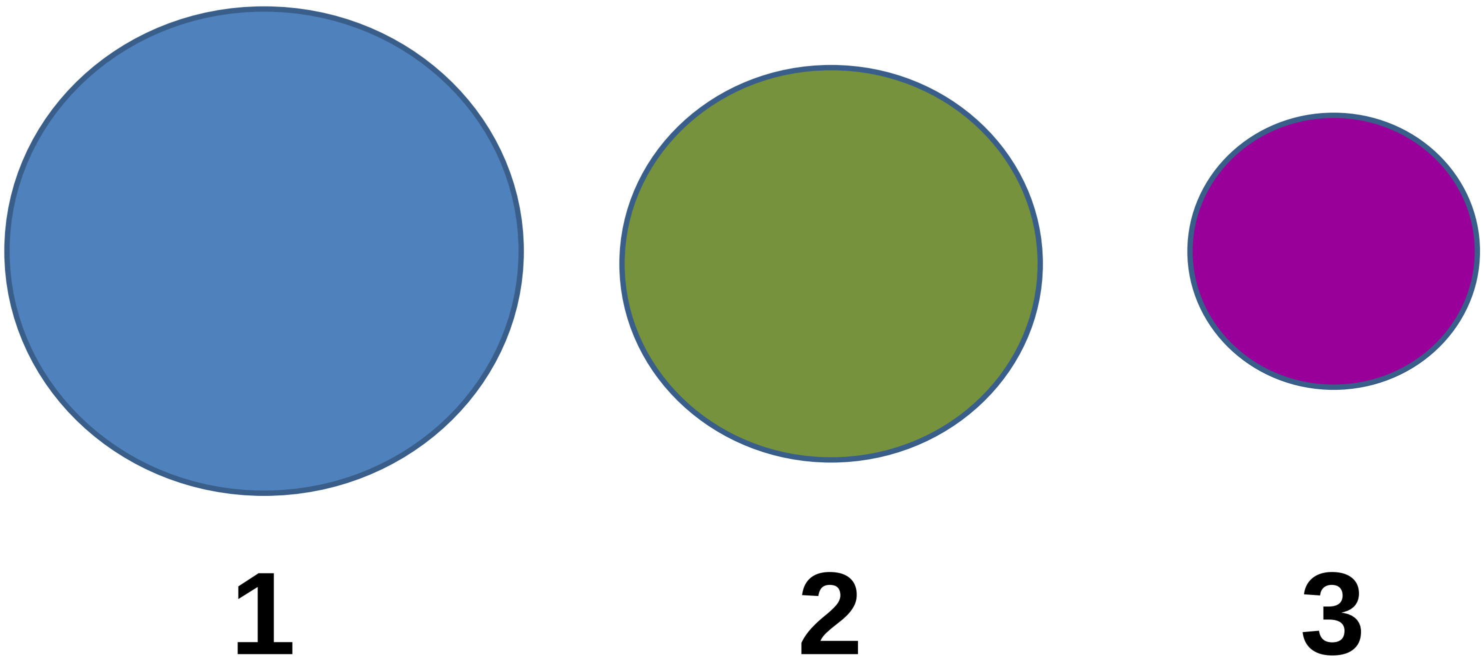
- Anion is formed when an atom accept electron.
- Neutral atom (Cl) & anion (Cl^-) have same no. of proton, but Cl^- ion contains more electrons than its neutral atom.
- Therefore, for Cl^- ion the mutual repulsion between the electrons increases.
- The repulsion caused the electron cloud spread out and the outer orbital expand.
- Thus, the size of anion is larger than its neutral atom.

Comparison Atomic Radii & Its Ionic Radii

cations (+ve ions)	anions (-ve ions)
formed when a neutral atom loses electron(s)	formed when a neutral atom gains electron(s)
less electrons repulsion : ~ electrons are removed from the valence shell	greater electrons repulsion : ~ electrons are added to the valence shell
remaining electrons are pulled closer to the nucleus	electrons are spread out
stronger attraction between nucleus & electrons	weaker attraction between nucleus & electrons
electron cloud shrink	electron cloud expand
size cation < neutral atom	size anion > neutral atom

Visualising Concept

Which of these spheres represent F, which represent Br, and which represent Br^{-} ?



Isoelectronic

- Isoelectronic species is a group of atoms or ions having the **same electronic configuration.**

- Example...

Ions	Electronic Configuration
Na^{+}	$1s^2\ 2s^2\ 2p^6$
Mg^{2+}	$1s^2\ 2s^2\ 2p^6$
Al^{3+}	$1s^2\ 2s^2\ 2p^6$
N^{3-}	$1s^2\ 2s^2\ 2p^6$
O^{2-}	$1s^2\ 2s^2\ 2p^6$
F^{-}	$1s^2\ 2s^2\ 2p^6$

- Na^{+}, Mg^{2+}, Al^{3+} and N^{3-}, O^{2-}, and F^{-} are isoelectronic ions with the electronic configuration as $1s^2\ 2s^2\ 2p^6$

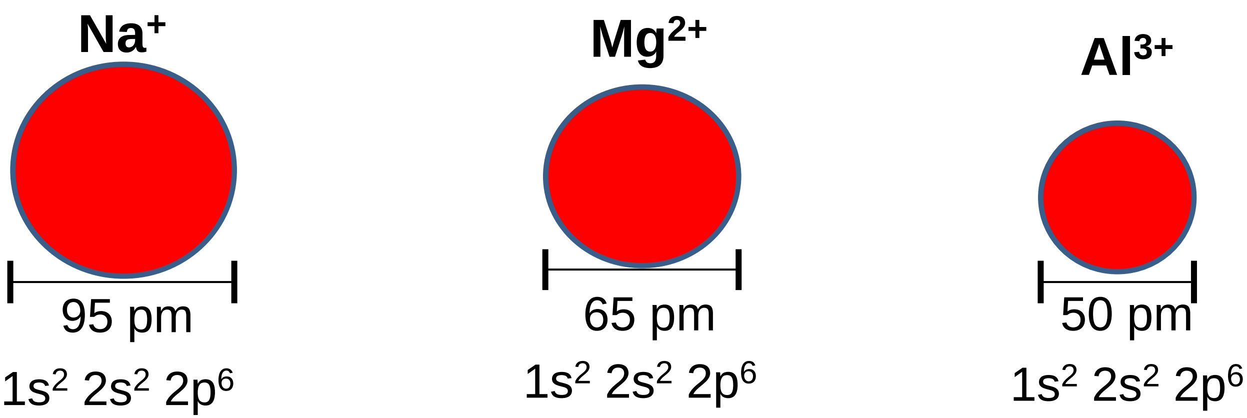
WHY?

Isoelectronic

Within isoelectronic series,

- The more positive the charge, the smaller the species.

Period 3



- The more negative the charge, the larger the species.

Period 2

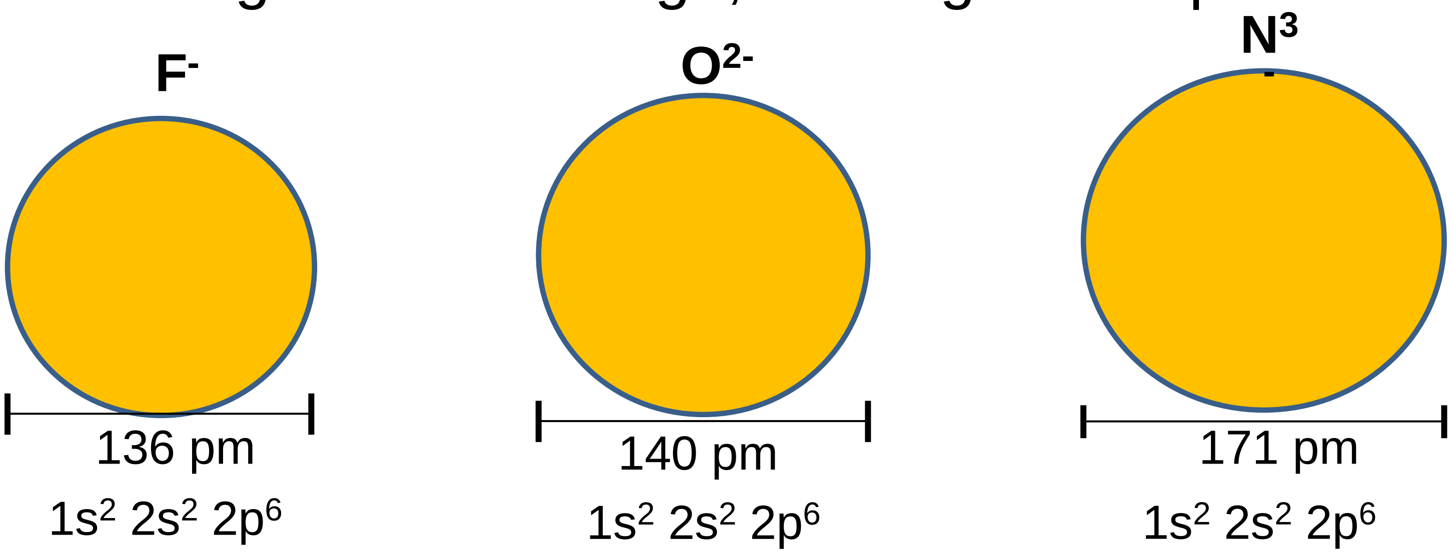


Figure show ionic radii of isoeletronic species (period 2 & 3)

Reason……

Isoelectronic

Ions	Electronic Configuration	Proton number	Z_{eff}
Al^{3+}	$1s^2\ 2s^2\ 2p^6$	13	+11
Mg^{2+}	$1s^2\ 2s^2\ 2p^6$	12	+10
Na^{+}	$1s^2\ 2s^2\ 2p^6$	11	+9
F^{-}	$1s^2\ 2s^2\ 2p^6$	9	+7
O^{2-}	$1s^2\ 2s^2\ 2p^6$	8	+6
N^{3-}	$1s^2\ 2s^2\ 2p^6$	7	+5

- ✓ When proton number increase, the effective nuclear charge also increase.
- ✓ The attraction between nucleus and remaining electron stronger.
- ✓ Thus,the ionic radii decrease.

Therefore the ionic radius $N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$

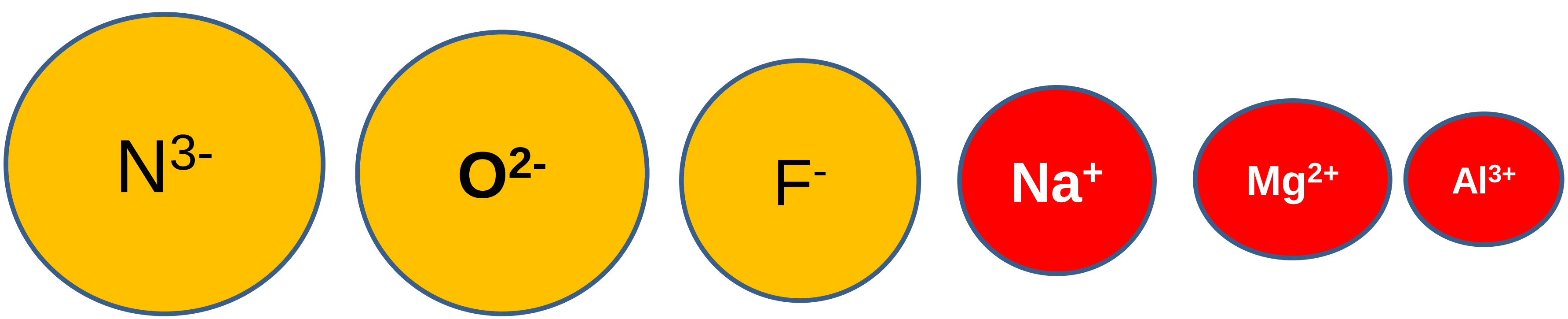
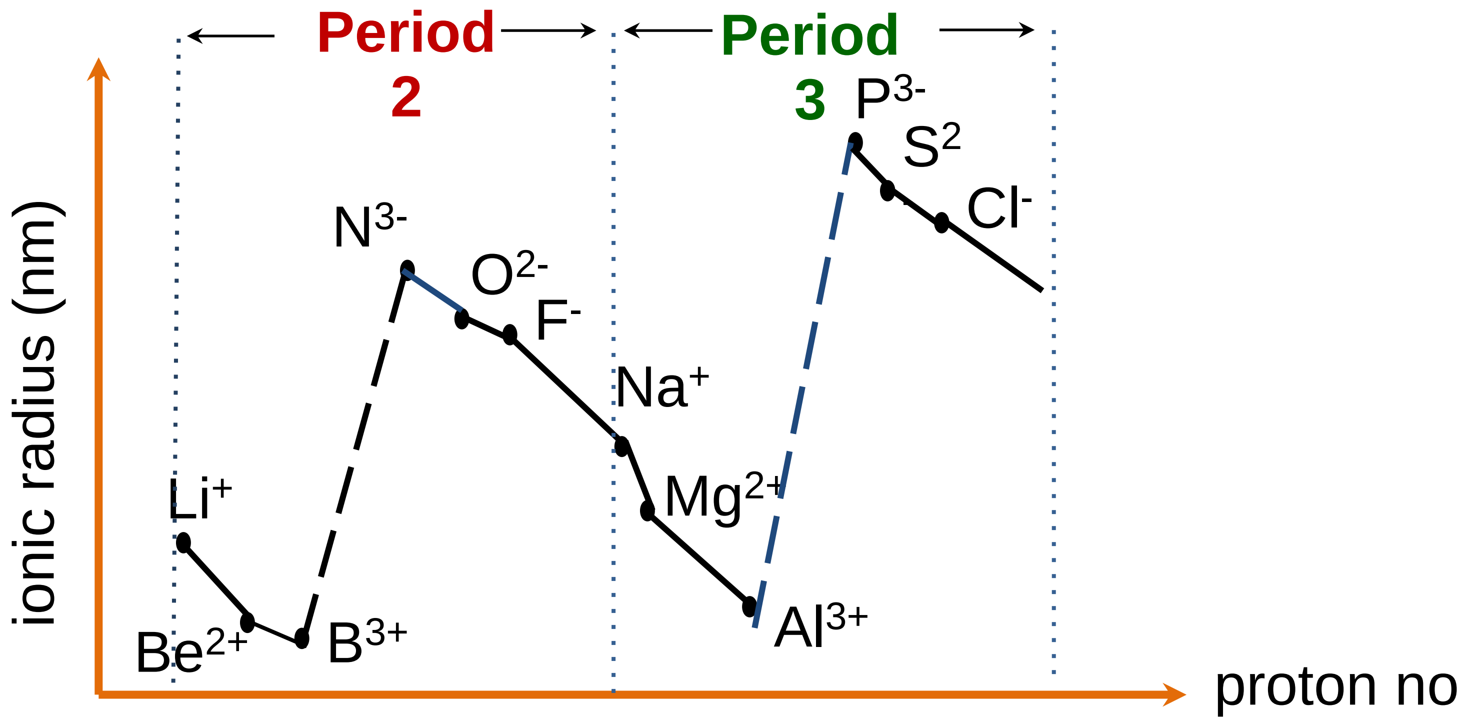


Figure show the ionic decreases as the effective nuclear charge increases.

Variation in the Ionic Radii Across Period 2 & 3



for isoelectronic species radii decrease from:

- Li^{+} to B^{3+}
- N^{3-} to F^{-}
- Na^{+} to Al^{3+}
- P^{3-} to Cl^{-}

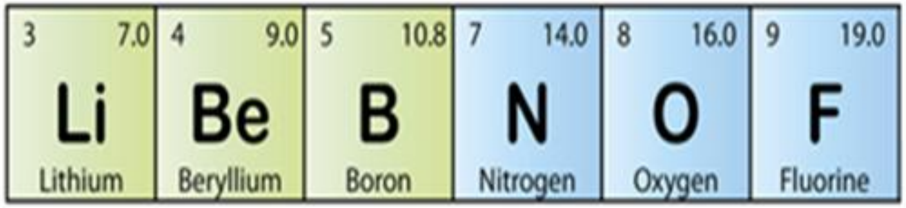
BECAUSE

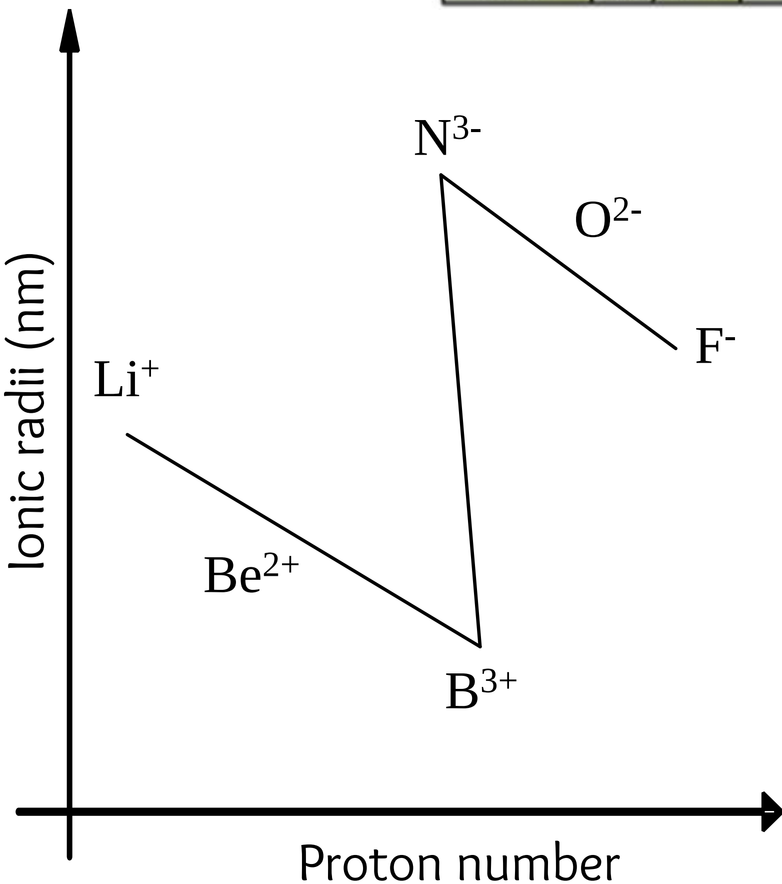
proton number increases

effective nuclear charge increases

Nucleus electron attraction stronger

Variation in the Ionic Radii Across Period 2





Across the period 2,

- ionic radii of cations decrease from Li^{+} to B^{3+}
- ionic radii of anions decrease from N^{3-} to F^{-}
- But the ionic radii increase drastically from B^{3+} to N^{3-}

WHY??!!

Variation in the Ionic Radii Across Period 2

1. Ionic radii decrease from Li^+ to B^{3+} because number of electrons is the same (isoelectronic species).

- the effective nuclear charge increases,
- the greater attraction between nucleus and the outer electrons,
- the size decreases.

$$Li^+ > Be^{2+} > B^{3+}$$

2. There is a large increase in ionic radii from B^{3+} to N^{3-} ion:

- N^{3-} has more shells (the higher value of n) and weaker attraction between nucleus and the outer electrons.
- thus, size of N^{3-} increases.

$$N^{3-} > B^{3+}$$

3. Ionic radii decrease from N^{3-} to F^{-} because number of electrons is the same, 10 electrons (isoelectronic species).

- the effective nuclear charge increases,
- the greater attraction between nucleus and the outer electrons,
- the size decreases.

$$N^{3-} > O^{2-} > F^{-}$$

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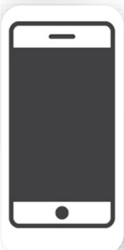
Application of Lithium-Ion Batteries

Consumer & Electronic devices

Li-Ion

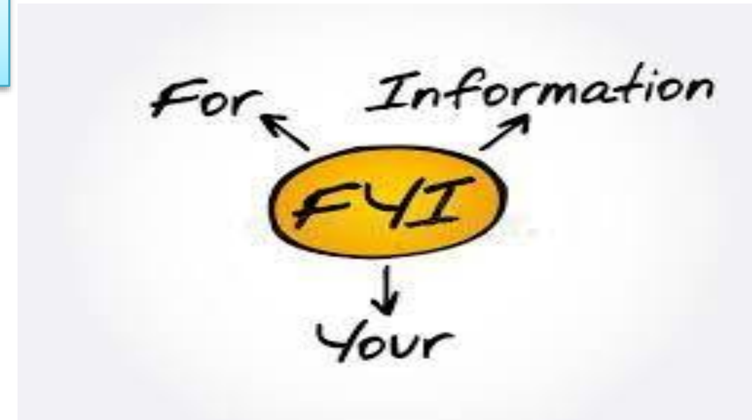
Transportation



https://www.istockphoto.com/

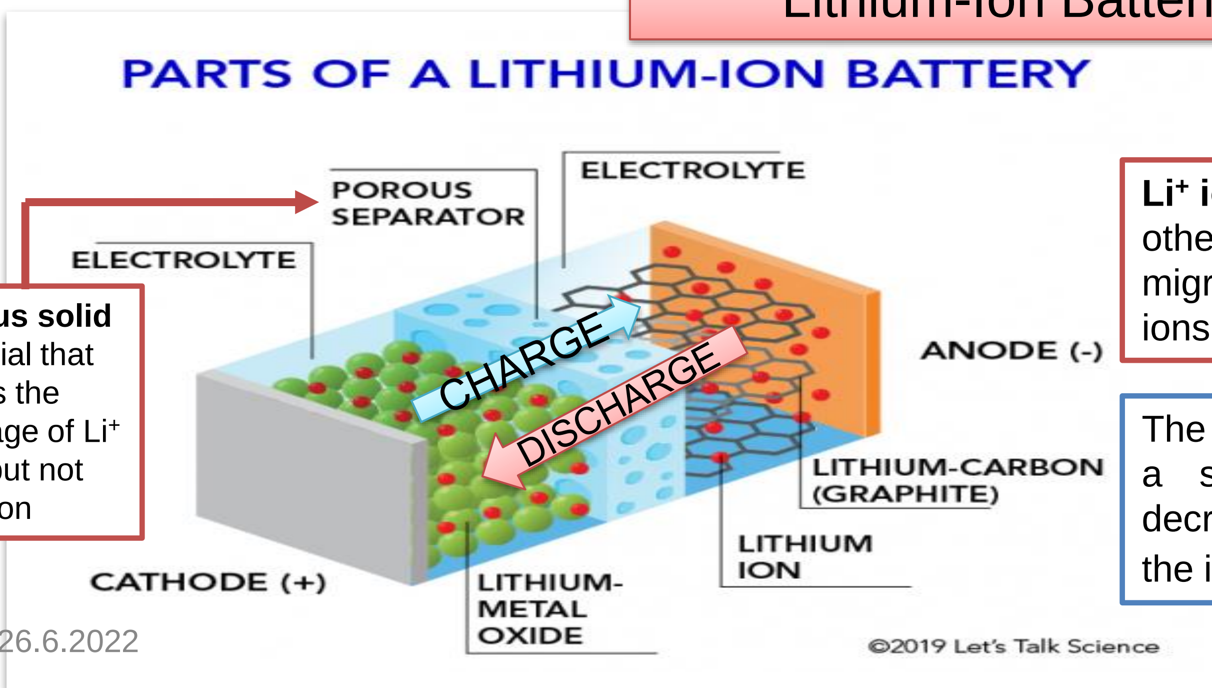
Energy storage

https://www.istockphoto.com/



Advantage:
- Rechargable
- light weight

Lithium-Ion Batteries & Ionic Size



Porous solid material that allows the passage of Li^+ ions but not electron

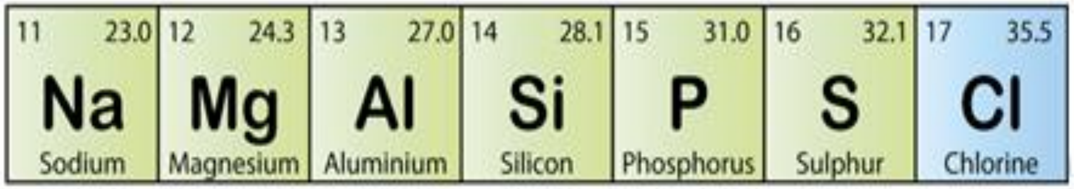
Li^+ ions are **smaller** than most other cations (allows them to migrate more readily than other ions can)

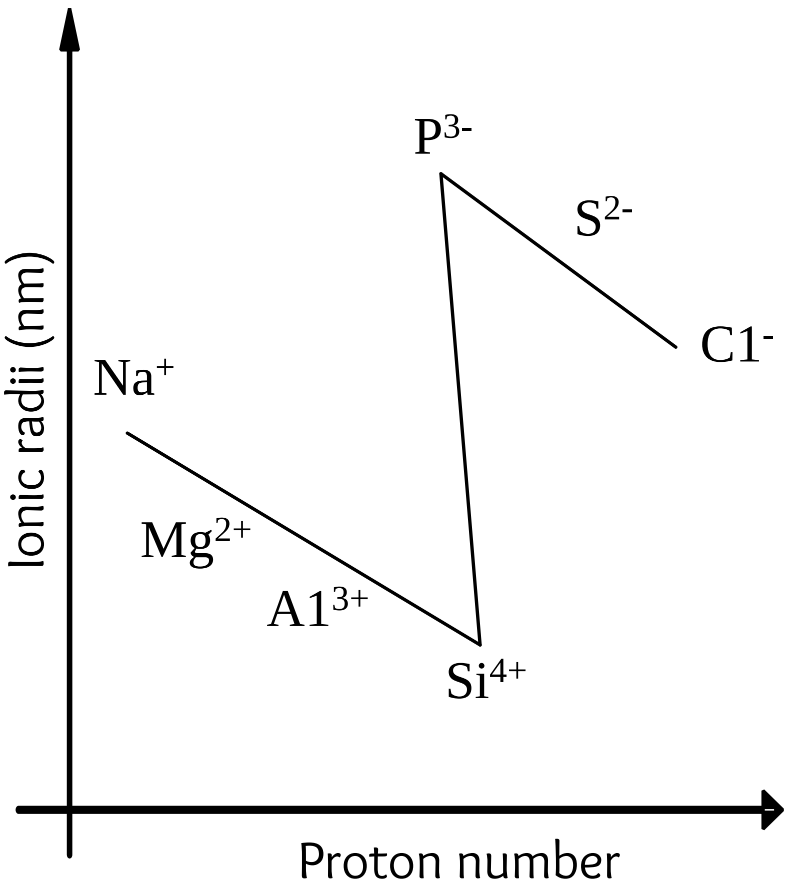
The ability of ion to move through a solid increase as the ion decreases and as the charge of the ions decreases.

Variation in the Ionic Radii Across Period 3

Ions	Electronic Configuration	Num of é	Z_{eff}	Ionic Radii (pm)
$_{11}Na^{+}$	$1s^2\ 2s^2\ 2p^6$	10	11 - 2 = +9	95
$_{12}Mg^{2+}$	$1s^2\ 2s^2\ 2p^6$	10	12 - 2 = +10	65
$_{13}Al^{3+}$	$1s^2\ 2s^2\ 2p^6$	10	13 - 2 = +11	50
$_{14}Si^{4+}$	$1s^2\ 2s^2\ 2p^6$	10	14 - 2 = +12	41
$_{15}P^{3-}$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	18	15 - 10 = +5	212
$_{16}S^{2-}$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	18	16 - 10 = +6	184
$_{17}Cl^{-}$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	18	17 - 10 = +7	181

Variation in the Ionic Radii Across Period 3





Across the period 3,

- ionic radii of cations decrease from Na^+ to Si^{4+}
- ionic radii of anions decrease from P^{3-} to Cl^-
- But the ionic radii increase drastically from Si^{4+} to P^{3-}

WHY??!!

Variation in the Ionic Radii Across Period 3

1. Ionic radii decrease from Na^+ to Si^{4+} because number of electrons is the same (isoelectronic species).

- the effective nuclear charge increases,
- the greater attraction between nucleus and the outer electrons,
- the size decreases.

$Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$

2. There is a large increase in ionic radii from Si^{4+} to P^{3-} ion:

- P^{3-} has more shells (the higher value of n) and weaker attraction between nucleus and the outer electrons.
- thus, size of P^{3-} increases.

$P^{3-} > Si^{4+}$

3. Ionic radii decrease from P^{3-} to Cl^{-} because number of electrons is the same, 18 electrons (isoelectronic species).

- the effective nuclear charge increases,
- the greater attraction between nucleus and the outer electrons,
- the size decreases. **$P^{3-} > S^{2-} > Cl^{-}$**

Ionisation Energy

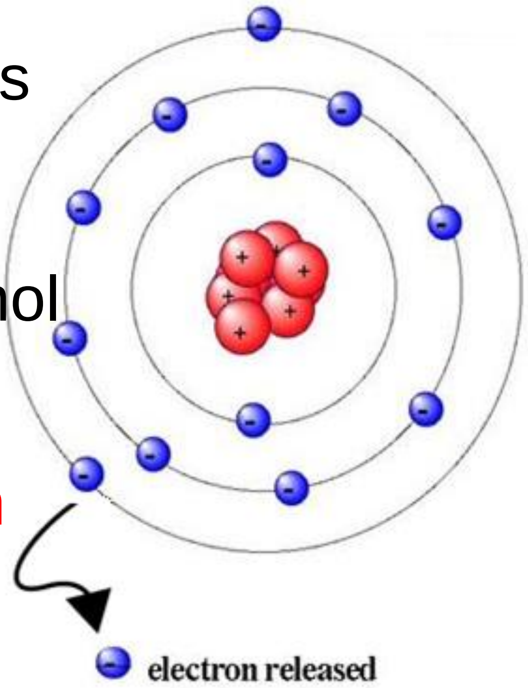
Ionisation energy is the minimum energy required to remove an electron from a gaseous atom or ion in its ground state.

✓ The First Ionisation Energy

- Is the minimum energy (in kJ/mol) required to remove an electron from a gaseous atom in its ground state.

Example :

$Mg_{(g)} \rightarrow Mg^{+}_{(g)} + e^{-}$ $\Delta H_1 = +736$ kJ/mol



✓ The Second Ionisation Energy

- Is the minimum energy required to remove an electron from positive gaseous ion in its ground state

Example :

$Mg^{+}_{(g)} \rightarrow Mg^{2+}_{(g)} + e^{-}$ $\Delta H_2 = +1450$ kJ/mol

Ionisation Energy

Factors Affecting the Ionization Energy, IE

Atomic radius	Effective nuclear charge, Z_{eff}	Shielding effect
✓ atomic radius increases, ✓ Ionization Energy decreases	✓ effective nuclear charge increases, the attraction between valence electrons and the nucleus becomes stronger ✓ Ionization Energy increases	✓ shielding effect increases, the attraction between valence electrons and the nucleus becomes weaker ✓ Ionization Energy decreases

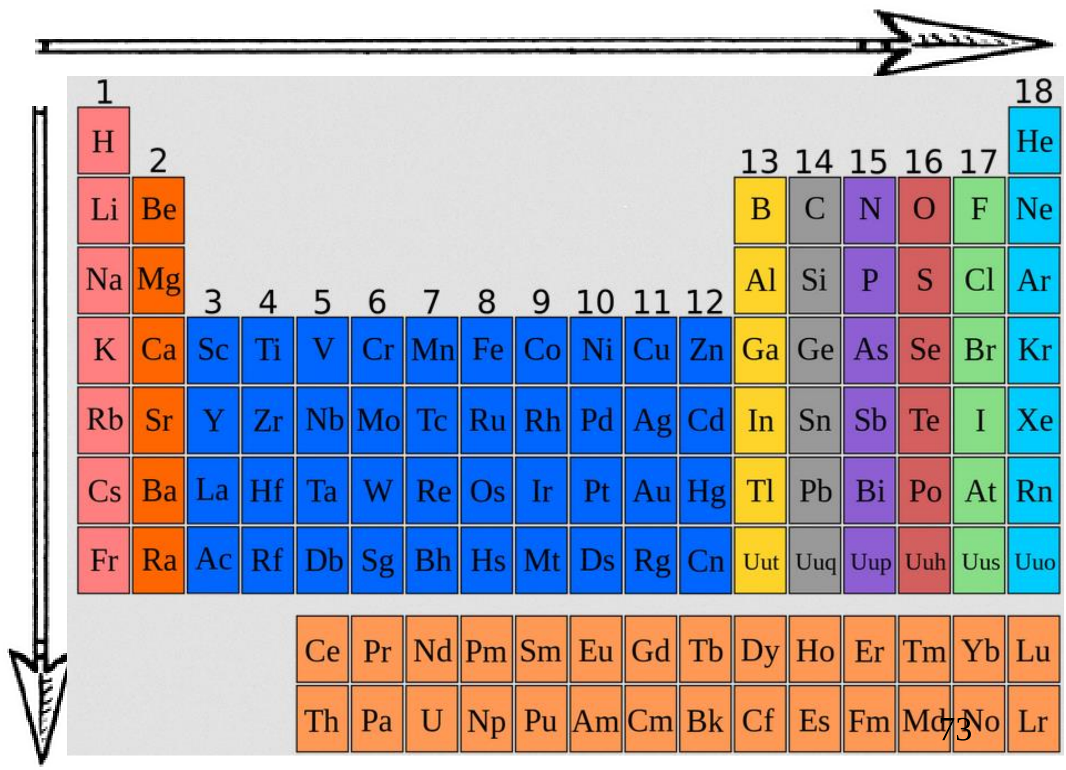
Ionisation Energy

Across a period :

- proton number increases
- effective nuclear charge, Z_{eff} increases
- attraction between valence electrons and nucleus become stronger
- ionisation energy increases

Down a group :

- no of shell, *n* increases
- shielding effect increases.
- attraction between valence electrons and nucleus become weaker.
- ionisation energy decreases



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Example 1

Arrange the followings increasing order of ionization energy.

$_{4}Be$, $_{12}Mg$ and $_{20}Ca$

Element	Electronic Configuration	Valence electronic configuration	Group
Be	$1s^2\ 2s^2$	$2s^2$	2
Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$	$3s^2$	2
Ca	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	$3s^2$	2

The elements belong to the same group

Answer:

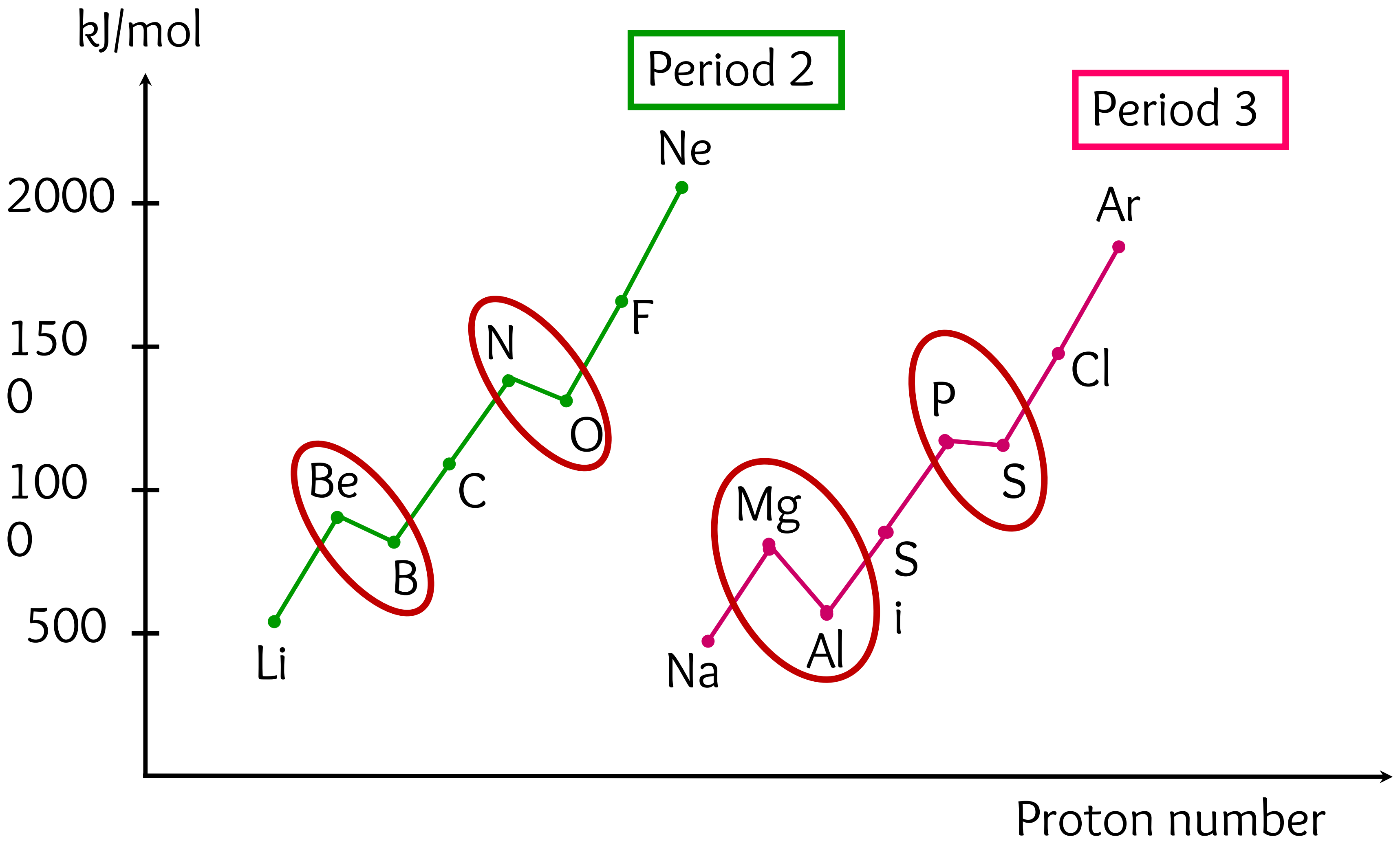
Increasing order of IE: Ca < Mg < Be

-The first ionization energy decreases on going down a group as the atomic radius increase in the order Be < Mg < Ca.

-Shielding effect increases, the attraction between valence electrons and the nucleus becomes weaker.

Ionisation Energy

Anomalous Cases for the First Ionisation Energy Across Period 2 and 3

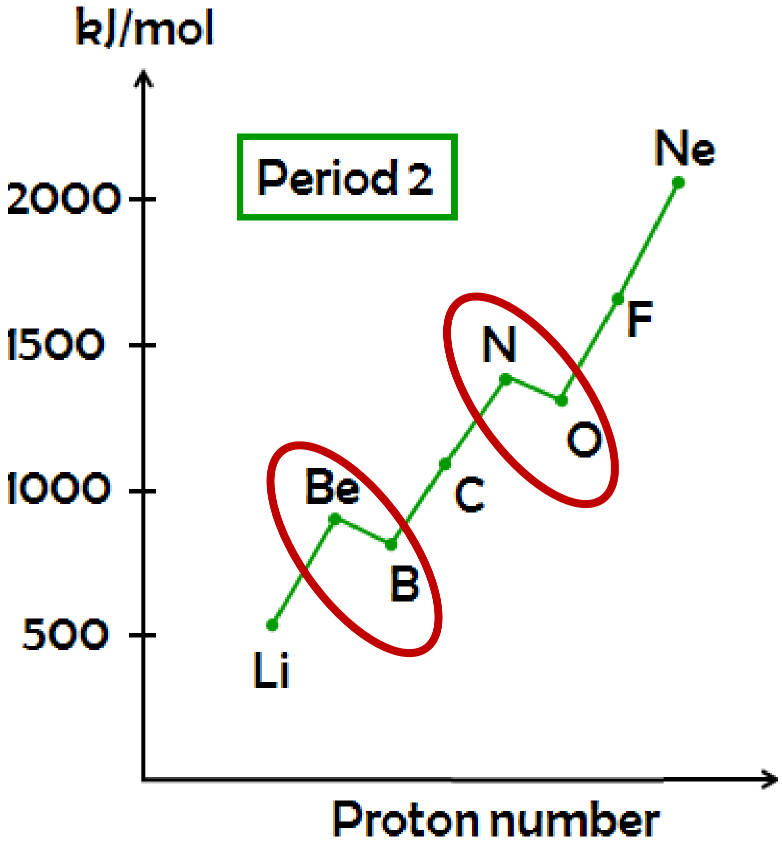


Ionisation Energy

Anomalous Cases for the 1st IE Across Period 2

- The increase of IE in proton number is not uniform.

- In Period 2, there are 2 cases:
  i. Group 2 and 13 : Be and B
  ii. Group 15 and 16 : N and O



Ionisation Energy

Anomalous Cases for Be and B (Period 2)

✓ Between Be (Group 2) and B (Group 13),

Electronic configuration: Be : $1s^2\ 2s^2$

B : $1s^2\ 2s^2\ 2p^1$

- 2p orbital of B is well shielded by the inner and the 2s electron.
- attraction between the nucleus and 2p electron is weak.
- less energy is needed to remove the electron in 2p orbital.

✓ Therefore, ionization energy of Be > B.

Anomalous Cases for N and O (Period 2)

✓ Between N (Group 15) and O (Group 16),

Electronic configuration:

N : $1s^2\ 2s^2\ 2p^3$ (**half-filled 2p orbital**)

O : $1s^2\ 2s^2\ 2p^4$ (**partially-filled 2p orbital**)

- Half filled 2p orbital of N is more stable than partially filled orbital of O
- More energy is needed to remove the electrons at the more stable orbital

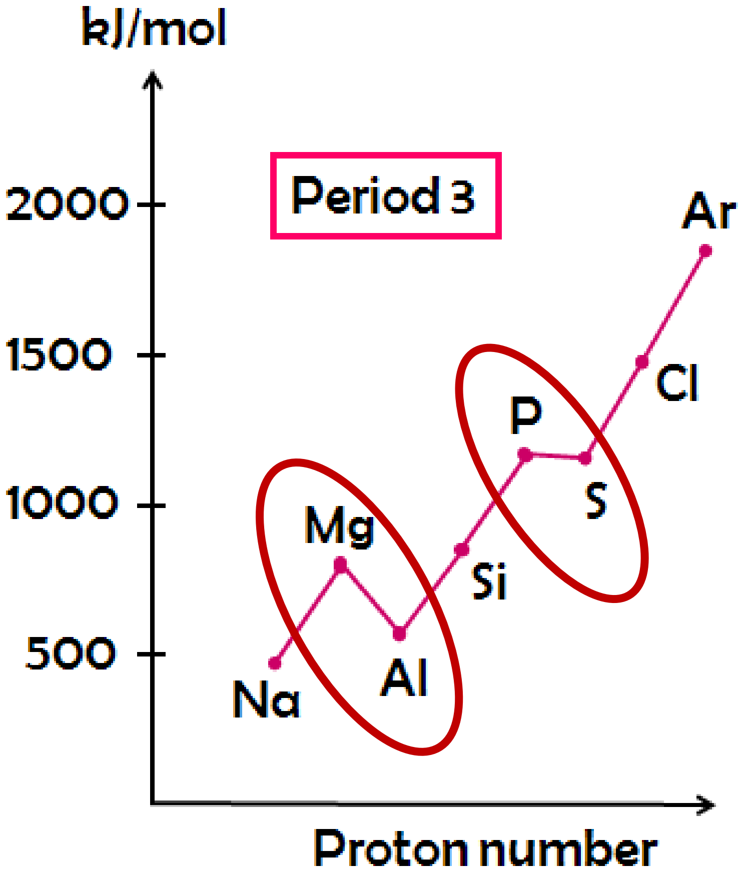
✓ Therefore, ionization energy of N > O

Anomalous Cases for the 1st IE Across Period 3

- The increase of IE in proton number is not uniform.
- In Period 3, there are 2 cases:

i. Group 2 and 13 : Mg and Al

ii. Group 15 and 16 : P and S



Anomalous Cases for Mg and Al (Period 3)

✓ Between Mg (Group 2) and Al (Group 13),

Electronic configuration: Mg : $1s^2\ 2s^2\ 2p^6\ 3s^2$

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

- 3p orbital of **Al** is well shielded by the inner and the 3s electrons
- attraction between the nucleus and 3p electron is weak.
- less energy is needed to remove the electron in 3p orbital.

✓ Therefore, ionization energy of Mg > Al.

Ionisation Energy

Anomalous Cases for P and S (Period 3)

✓ Between P (Group 15) and S (Group 16),

Electronic configuration:

P : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$ **(half-filled 3p orbital)**

S : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ **(partially-filled 3p orbital)**

- Half filled 3p orbital of P is more stable than partially filled 3p orbital of S
- More energy is needed to remove the electrons at the more stable orbital

✓ Therefore, ionization energy of P > S.

Successive Ionisation Energy

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

Silberberg, Pg.331

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

	first	second	third	fourth	fifth	sixth	seventh
Li	520	7297	11810	-	-	-	-
Be	900	1757	14840	21000	-	-	-
B	800	2430	3659	25020	32810	-	-

Successive Ionisation Energy

Example of Be atom

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

1st e⁻ removed from 2s $1s^2\ 2s^2$ $1s^2\ 2s^1$

$Be^{+}(g) \rightarrow Be^{2+}(g) + e^{-}$ **IE_2= 1757 kJ/mol**

2nd e⁻ removed from 2s $1s^2\ 2s^1$ $1s^2$

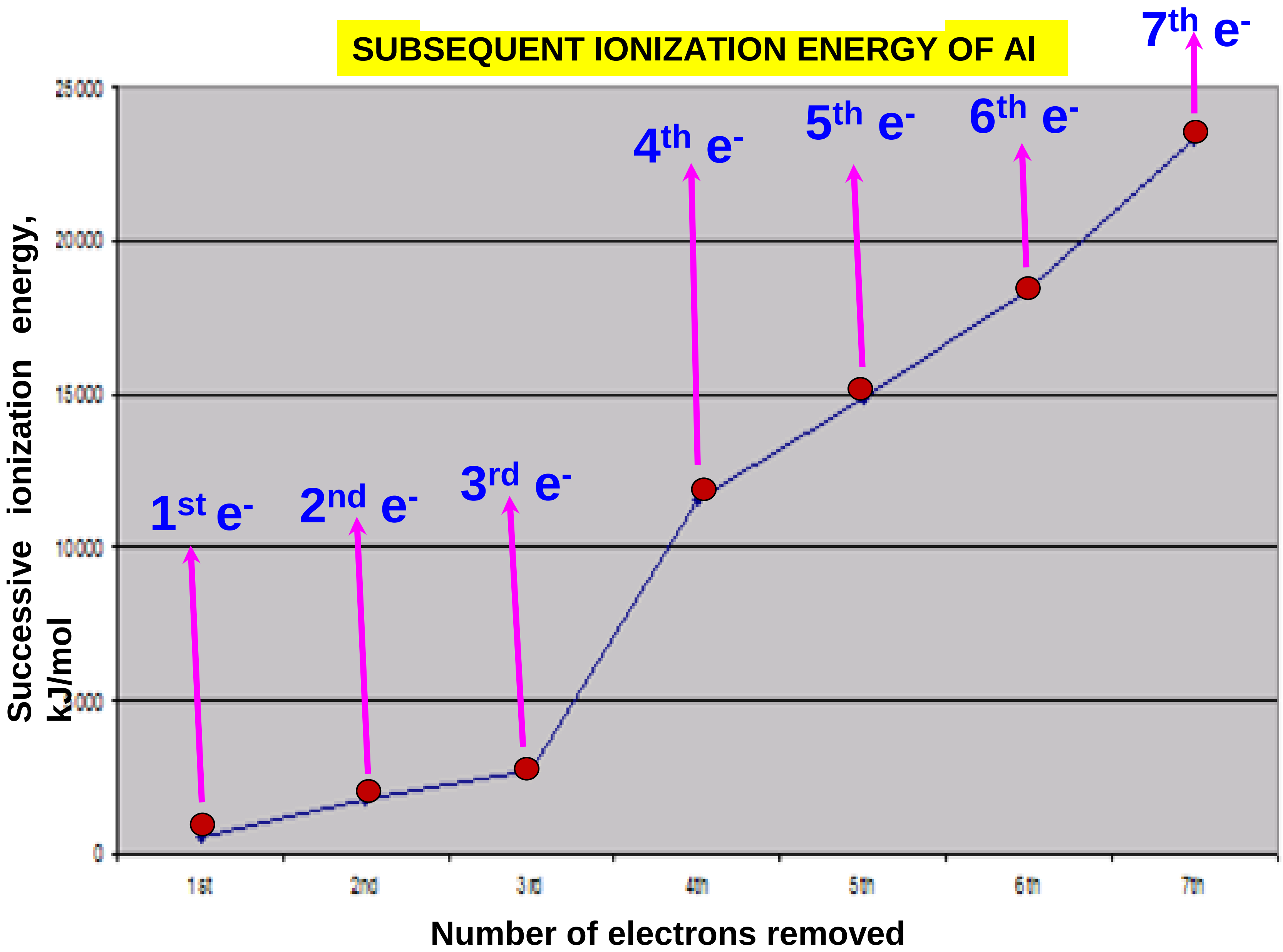
$Be^{2+}(g) \rightarrow Be^{3+}(g) + e^{-}$ **IE_3= 14840 kJ/mol**

3rd e⁻ removed from 1s $1s^2$ $1s^1$

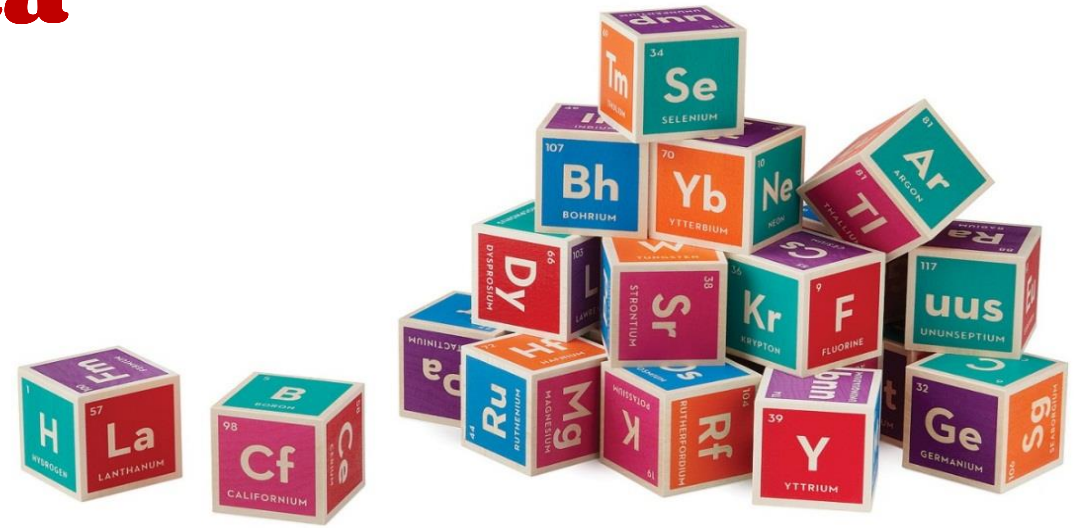
$Be^{3+}(g) \rightarrow Be^{4+}(g) + e^{-}$ **IE_4= 21000 kJ/mol**

4th e⁻ removed from 1s $1s^1$

Successive Ionisation Energy Graph



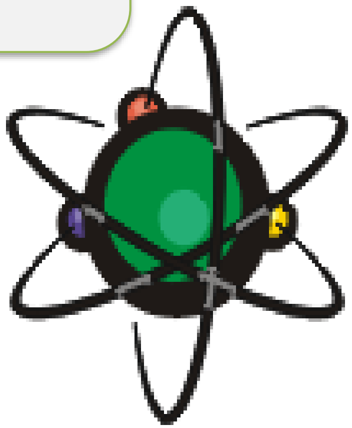
Deduce the Electronic Configuration of an Element and its Position in the Periodic Table Based on Successive Ionisation Energy Data



Successive Ionisation Energy

From the data of successive ionisation energy, we can deduce:

- ✓ Number of valence electron
- ✓ Group number of the element
- ✓ Electronic configuration



Successive Ionisation Energy

By analysing the data of successive ionisation energy,

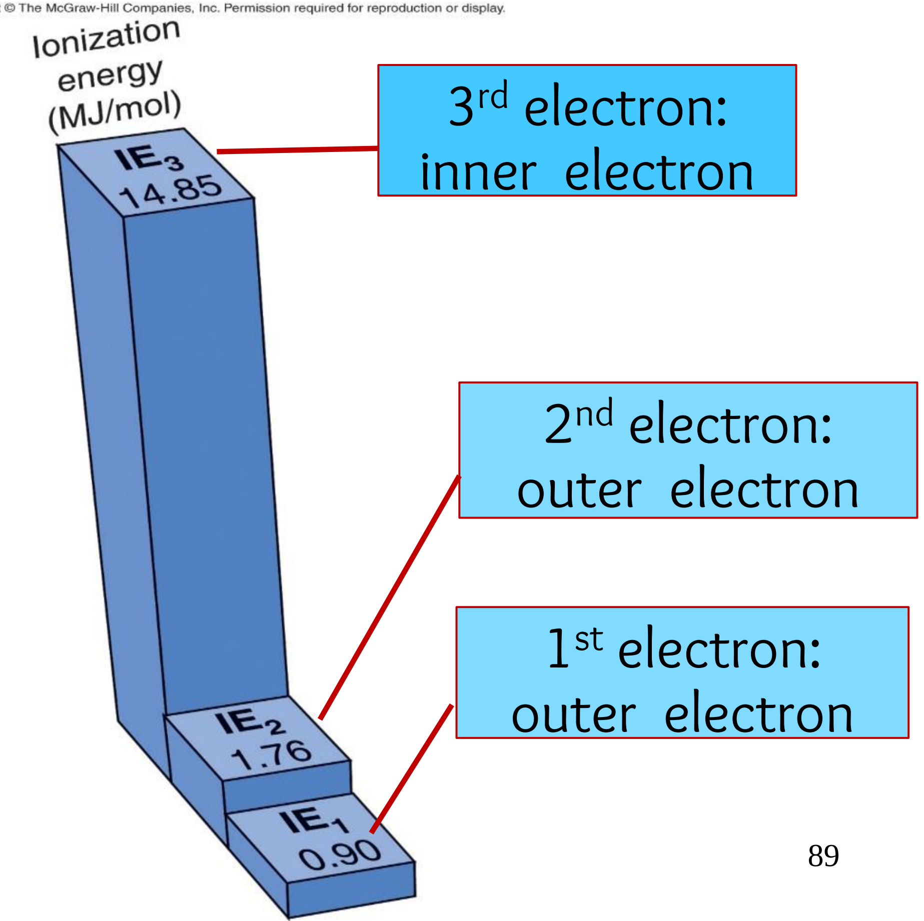
- enable to figure out the position of element in the Periodic Table by looking for the highest ratio @ sudden increase.

Element	Electronic configuration	First	Second	Third	Fourth	Fifth
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

Successive Ionisation Energy

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.

Example : Be (Z = 4) :1s^2 2s^2



Element	Electronic configuration	First	Second	Third	Fourth	Sixth	Seven
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** : IE_2 to IE_1 have the highest ratio @ sudden increase is between IE_1 to IE_2 because the second electron is removed from *1s* orbital (inner shell).

For **Be** : IE_3 to IE_2 have the highest ratio @ sudden increase is between IE_2 to IE_3 because the third electron is removed from *1s* orbital (inner shell).

For **B** : IE_4 to IE_3 have the highest ratio @ sudden increase is between IE_3 to IE_4 because the fourth electron is removed from *1s* orbital (inner shell).

Successive Ionisation Energy

- Therefore, we can determine the electronic configuration of the valence electron for an element using following methods:

  - Method 1:-

    By determining the IE ratios.

  - Method 2:-

    By determining the differences in IE.

Successive Ionisation Energy

Based on the information given below, determine the group of the element. Explain.

Element	Electronic configuration	First IE 1	Second IE 2	Third IE 3	Fourth IE 4		
Be	$1s^2\ 2s^2$	**900**	**1760**	**14850**	**21910**	-	-

- Method 1:-

By determining the IE ratios:

$$\frac{IE2}{IE1} = \frac{1760}{900} = 1.96$$

$$\frac{IE3}{IE2} = \frac{14850}{1760} = \mathbf{8.44}$$

$$\frac{IE4}{IE3} = \frac{21910}{14850} = 1.48$$

Successive Ionisation Energy

- The first and second electron are removed from the same energy subshell (2s).

- The third electron is removed from an inner shell that is 1s, hence it requires a higher IE3 (**8.44 times**) than IE2.

- Since IE3 / IE2 have the highest ratio, **2 valence electrons** are present.

- Electronic configuration: **ns^2**
- This element is in **Group 2**

Successive Ionisation Energy

- Method 2:-

By determining the differences in IE:

IE2 - IE1 = 860 kJ mol^{-1}

IE3 - IE2 = **13090** kJ mol^{-1}

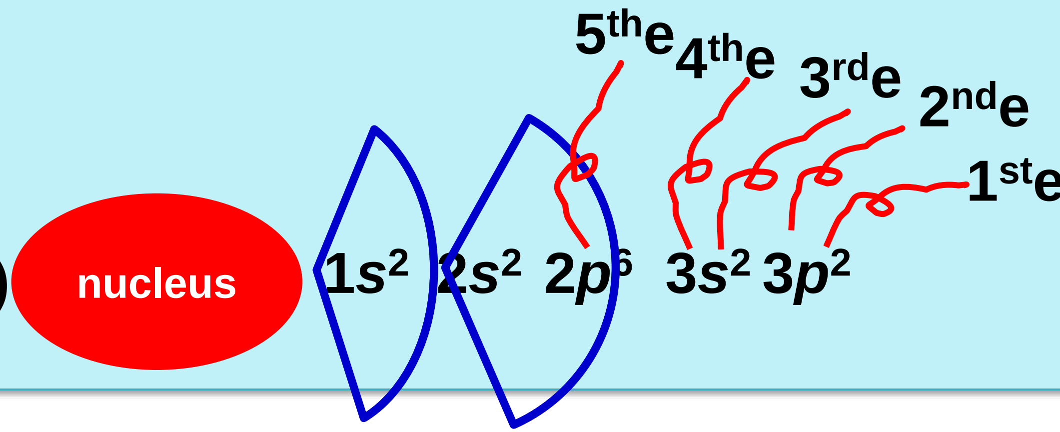
IE4 - IE3 = 7060 kJ mol^{-1}

Successive Ionisation Energy

- The first and second electron are removed from the same energy subshell (2s).
- The third electron is removed from an inner shell that is 1s, hence it requires a higher IE3 than IE2 (a difference of **13090** kJ mol-1).
- Since IE3 - IE2 have the highest difference, 2 valence electrons are present.
- Electronic configuration: ns^2
- This element is in Group 2

EXAMPLE:

Si (Z = 14)

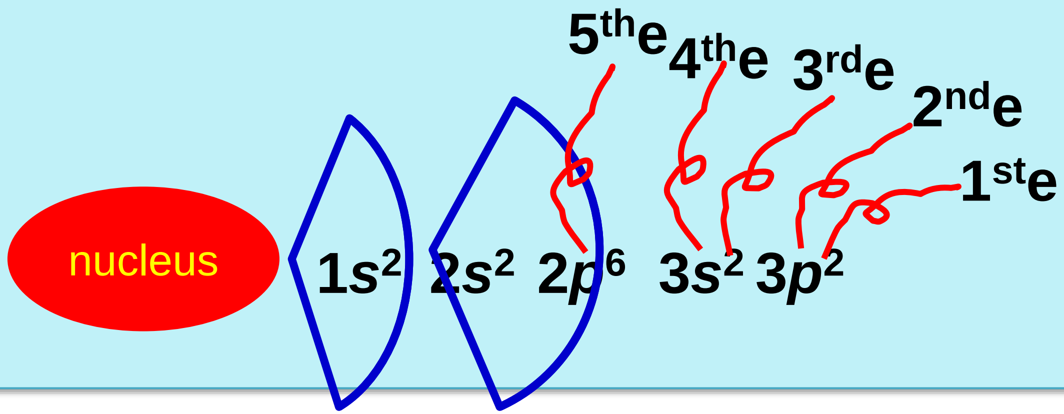


786 kJ/mol	IE_1: 1^{st} e^- removed from ***3p*** subshell
1580 kJ/mol	IE_2: 2^{nd} e^- removed from ***3p*** subshell
3230 kJ/mol	IE_3: 3^{rd} e^- removed from ***3s*** subshell
4360 kJ/mol	IE_4: 4^{th} e^- removed from ***3s*** subshell
16100 kJ/mol	IE_5: 5^{th} e^- removed from ***2p*** subshell,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

Highest ratio (IE5/IE4 = 3.7)/ sudden increase

EXAMPLE:

Si (Z = 14)



- IE_5 to IE_4 have the highest ratio (3.7) @ drastic increase in ionisation energy from IE_4 to IE_5
- The 5^{th} 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 = **4**   Group = **14**
- Valence electron configuration = $\boldsymbol{3s^2 3p^2}$

Example 1 :

Table shows all four successive ionisation energies of X. *Deduce* the valence electrons, group no. & valence electronic configuration from the data of successive ionisation energies

IE	1	2	3	4
$kJ\ mol^{-1}$	899	1757	14845	21000

ANSWER :

$$\frac{IE_2}{IE_1} = \frac{1757}{899} = 1.95$$

$$\frac{IE_4}{IE_3} = \frac{21000}{14845} = 1.41$$

$$\frac{IE_3}{IE_2} = \frac{14845}{1757} = \mathbf{8.45}$$

Example 1 :

ANSWER :

- IE_3 to IE_2 have the highest ratio (8.45) @ drastic increase in ionisation energy from IE_2 to IE_3
- The 3^{rd} 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 = **2** Group = **2**
- Valence electron configuration = $\boldsymbol{ns^2}$

Example 2 :

Five successive ionization energies (kJ mol^{-1}) for atom M is shown below. Determine valence electrons, group no. & valence electronic configuration from the data of successive ionisation energies

IE_1	IE_2	IE_3	IE_4	IE_5
800	1580	3230	4360	16000

Determine :

(i) electronic configuration of the valence electron for M.

(ii) group number of M in the periodic table.

$$\frac{IE_2}{IE_1} = \frac{1580}{800} = 1.98$$

$$\frac{IE_4}{IE_3} = \frac{4360}{3230} = 1.35$$

$$\frac{IE_3}{IE_2} = \frac{3230}{1580} = 2.04$$

$$\frac{IE_5}{IE_4} = \frac{16000}{4360} = \mathbf{3.67}$$

Example 2 :

Five successive ionization energies (kJ mol^{-1}) for atom M is shown below. Determine valence electrons, group no. & valence electronic configuration from the data of successive ionisation energies

IE_1	IE_2	IE_3	IE_4	IE_5
800	1580	3230	4360	16000

- Ratio $\frac{IE_5}{IE_4}$ is the highest,
- there are 4 valence electrons,
- The fifth electron is removed from an inner shell.

(i) Electronic configuration of valence electron for M is $ns^2\ np^2$

(ii) M is in Group 14 at the periodic table

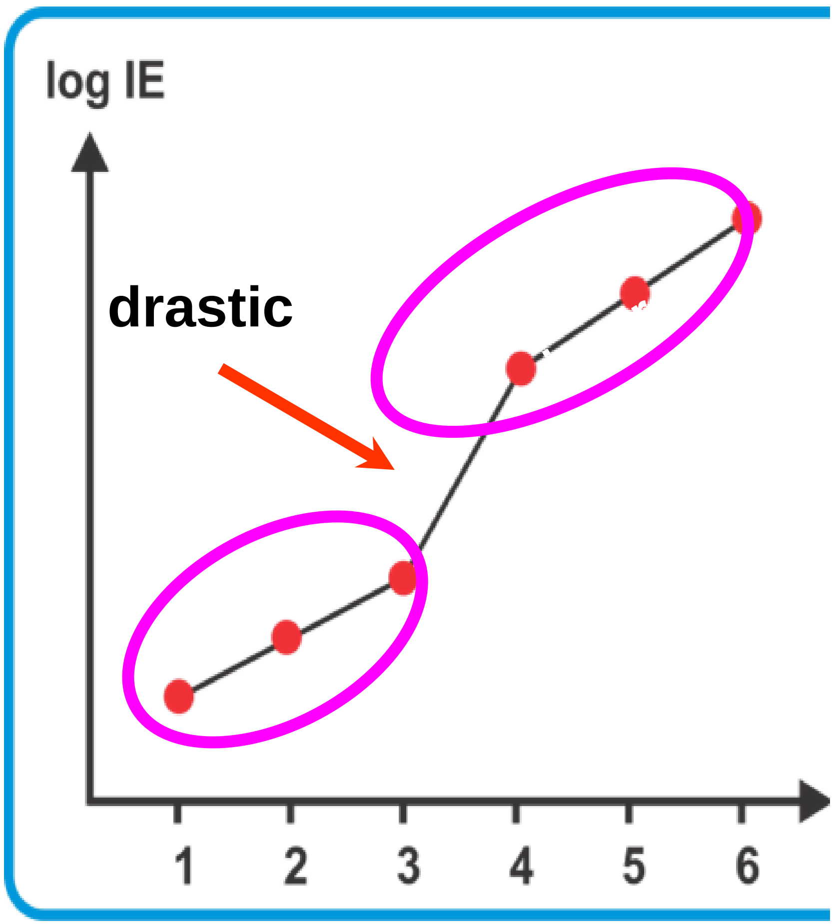
Example 3:

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

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

Example 3 :

Deduce the valence electrons, group no. & valence electronic configuration from the successive ionisation energies graph



- From the plot of successive IE, there is a sudden increase from IE_3 to IE_4.
- It means that more energy is required to remove the fourth electron.
- Since the fourth electron is difficult to be removed, the fourth electron is removed from an inner shell.
- Thus, there are 3 valence electrons
- Therefore, valence electronic configuration is $ns^2\ np^1$ and located in Group 13

Keep In Mind

The successive ionisation energies for an element is **only applicable** for main group elements

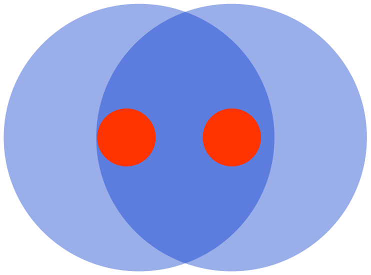
(block s & block p).

DEFINITION

Electronegativity is a relative tendency of an atom to attract electrons to itself when chemically combined with another atom.

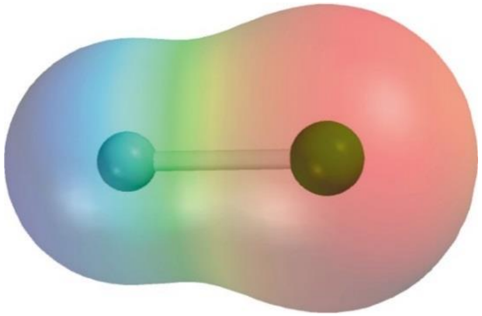
H : H

H : F



H_2

Non polar
covalent bond



HF

Polar
covalent bond

Common Element Electronegativities

Increasing electronegativity

Decreasing electronegativity

1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
H 2.1																	
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	
Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	
Fr 0.7	Ra 0.9																

Electronegativity

Down a group

- no of shell, *n* increases
- shielding effect increases
- attraction between valence electrons and nucleus weaker
- atomic size increases
- ability of an atom to attract the shared electrons decreases
- electronegativity decreases

Across a period

- proton number increases
- effective nuclear charge, Z_{eff} increases
- attraction between valence electrons and nucleus stronger
- atomic size decreases
- ability of an atom to attract the shared electrons increases
- electronegativity increases

Oxides of Element Period 3

Element	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
type of oxide	basic		amphoteric	acidic			
Type of compound	ionic			molecular			

Acid-Base Character of Oxides of Element in Period 3

Elements react with oxygen to produce oxides

For elements in period 3:

- ✓ $4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$
- ✓ $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
- ✓ $4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$

Acid-Base Character of Oxides of Element in Period 3

- ✓ $Si(s) + O_2(g) \longrightarrow SiO_2(s)$
- ✓ $P_4(s) + 5O_2(g) \longrightarrow P_4O_{10}(s)$
- ✓ $S(s) + O_2(g) \longrightarrow SO_2(g)$
- ✓ $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$
- ✓ $2Cl_2(g) + 7O_2(g) \longrightarrow 2Cl_2O_7(l)$

Basic Oxide

- ✓ Forms when Na or Mg reacts with oxygen.
- ✓ An ionic compounds.
- ✓ Na_2O react with water to form base NaOH

$$Na_2O(\boldsymbol{s}) + H_2O(\boldsymbol{l}) \rightarrow 2NaOH(\boldsymbol{aq})$$

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

$$MgO(\boldsymbol{s}) + 2HCl(\boldsymbol{aq}) \rightarrow MgCl_2(\boldsymbol{aq}) + H_2O(\boldsymbol{l})$$

Amphoteric Oxide

- ✓ Amphoteric : act as acid and base .
- ✓ Forms when Al reacts with oxygen.
- ✓ An ionic compounds.

- ✓ Al_2O_3 acts as acid :

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

- ✓ Al_2O_3 acts as base :

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

Acid-Base Character of Oxides of Element in Period 3

Acidic Oxide

- ✓ Forms when Si react with oxygen.
- ✓ An ionic compounds.
- ✓ SiO_2 is insoluble in water, however, it does react with bases to produce salt and water

$$SiO_2\,(\boldsymbol{s}) + 2NaOH\,(\boldsymbol{aq}) \rightarrow Na_2SiO_3(\boldsymbol{aq}) + H_2O\,(\boldsymbol{l})$$

Acid-Base Character of Oxides of Element in Period 3

Acidic Oxide

- ✓ Forms when P, S and Cl react with oxygen.
- ✓ A covalent compounds.
- ✓ Oxides react with water to produce acidic solution.

$$SO_3\,(\boldsymbol{g}) + H_2O\,(\boldsymbol{l}) \rightarrow H_2SO_4(\boldsymbol{aq})$$

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

$$Cl_2O_7\,(\boldsymbol{l}) + H_2O\,(\boldsymbol{l}) \rightarrow 2HClO_4(\boldsymbol{aq})$$

Acid-Base Character of Oxides of Element in Period 3

	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_6 or P_4O_{10}	SO_2 or SO_3	Cl_2O_7
adding H_2O	soluble	insoluble	insoluble	insoluble	soluble	soluble	soluble
adding HCl	soluble	soluble	soluble	insoluble	insoluble	insoluble	insoluble
adding NaOH	insoluble	insoluble	soluble	soluble	soluble	soluble	soluble
Character of oxides	basic	basic	amphoteric	acidic	acidic	acidic	acidic

Failure is success if we learn from it.
~ Malcolm Forbes

By:

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