

The image shows a periodic table with orbital subshells highlighted. The s-block is blue, the p-block is green, the d-block is red, and the f-block is orange. The table is labeled with atomic numbers 1 through 118.

1	2	10	16	18											
1s	s		p	1s											
2s		d	2p	Ne											
3s			3p	Ar											
4s		3d	4p	Kr											
5s		4d	5p	Xe											
6s		5d	6p	Rn											
7s		6d	7p												
f	La	Ce	Pr	Nd	Pm	Sm	Eu	4f	Tb	Dy	Ho	Er	Tm	Yb	Lu
	Ac	Th	Pa	U	Np	Pu	Am	5f	Bk	Cf	Es	Fm	Md	No	Lr

# CHAPTER 3: PERIODIC TABLE



# LEARNING OUTCOMES

- At the end of the lesson, students should be able to :
  - (a) Describe **period**, **group** and **block** (*s, p, d, f*). (C1)
  - (b) Deduce the **position of elements** in the periodic table from its electronic configuration. (C2)

# Period, Group and Block (*s,p,d,f*)

- The periodic table is a table that arranges all the known elements in order of increasing **proton number**.
- A **vertical** column of elements is called a **group**.
- A **horizontal** row of elements is known as a **period**.

# Groups

- The groups in the Periodic Table are numbered from 1 to 18.
- Elements in the same group have the same number of valence electrons.
- Example: Sodium and potassium are both found in group 1 which means that they both have 1 valence electron

Group number = number of valence electron

(if the element is in block s and d)

OR

Group number = number of valence electron + 10

(if the element is in block p)

1
H 1
Li 3
Na 11
K 19
Rb 37
Cs 55
Fr 87

# Periodic Table

Na 11	Mg 12	Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
----------	----------	----------	----------	---------	---------	----------	----------

Period 3

Group 1

1
H 1
Li 3
Na 11
K 19
Rb 37
Cs 55
Fr 87

1																		18	
1	H	2																	He
2	Li	Be																	Ne
3	Na	Mg	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
6	Cs	Ba	[ ]	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
7	Fr	Ra	[ ]	Rf	Db	Sg	Bh	Hs	Mt	Uun	Uuu	Uub		Uuq		Uuh		Uuo	
			[ ]											[ ]					
			La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu		
			Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr		

# Main group ( Representative Group ) in periodic table

- Group 1 : alkali metals
- Group 2 : alkaline earth metals
- Group 3-12: transition metals
- Group 16 : chalcogens
- Group 17 : halogens
- Group 18 : inert / noble gases

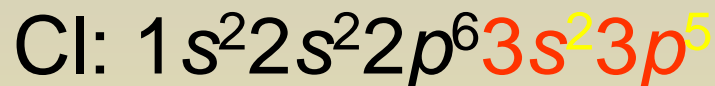
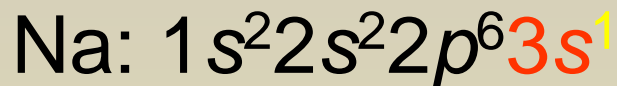


# Period

Na 11	Mg 12
----------	----------

Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
----------	----------	---------	---------	----------	----------

- The atoms of the elements in each period possess an identical number of **quantum levels**.
- Are numbered **1 to 7**.
- Example: Sodium and chlorine are in row 3 or **Period 3** because their principal quantum number of the valence electron shell is 3.



Period number = the largest principle quantum number, **n**, of the electrons in valence shell

# Blocks

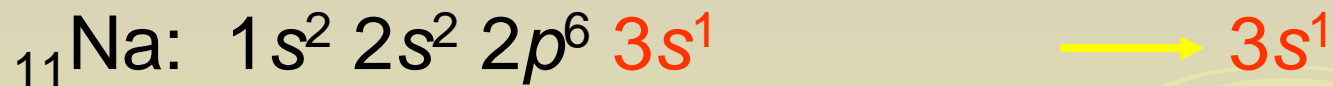
- All the elements in the periodic table can be classified into **4 main blocks** according to their valence electron configuration.
- These main blocks are block **s**, **p**, **d** and **f**.

1																			18
1	1s	2	s										13	14	15	16	17	1s	
2	Li	Be											3	4	5	6	7	8	9
3	Na	Mg												Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	3d	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	4d	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba		Hf	Ta	W	Re	5d	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra		Rf	Db	Sg	Bh	6d	Hs	Mt	Uun	Uuu	Uub		Uuq		Uuh		Uuo
			f																
	La	Ce	Pr	Nd	Pm	Sm	Eu	4f	Tb	Dy	Ho	Er	Tm	Yb	Lu				
	Ac	Th	Pa	U	Np	Pu	Am	5f	Bk	Cf	Es	Fm	Md	No	Lr				



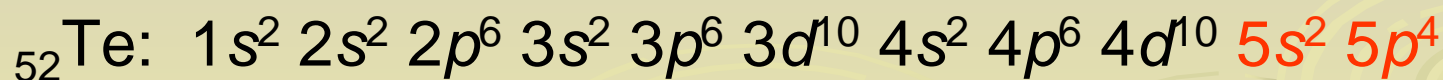
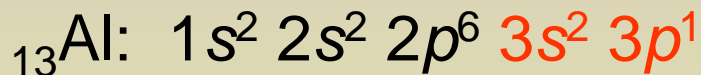
## s Block :

- Groups 1 and 2
- Filling of valence electron only involve s orbital
- Configuration of valence electron :  $ns^1$  to  $ns^2$
- Example:



## p Block :

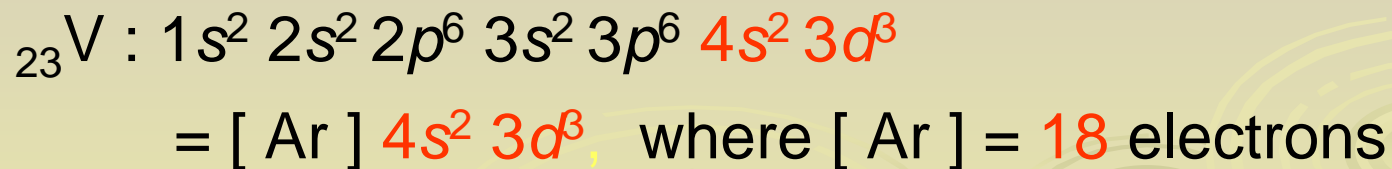
- Groups 13 to 18
- Configuration of valence electrons:  
 $ns^2 np^1$  to  $ns^2 np^6$ .
- Example:



## d Block

- Also known as a transition elements.
- Groups 3 to 12.
- Configuration of valence electron:  
 $ns^2 (n-1)d^1$  to  $ns^2 (n-1)d^{10}$

- Example:



## f Block

- Involve elements in the series of **lanthanides** (La to Lu) and **actinides** (Ac to Lr) in which the filling of valences electron happens in the subshell of **4f** and **5f**

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

Lanthanides

actinides

	1											18							
1	1s	2											13	14	15	16	17	1s	
2	Li	Be											B	C	N	O	F	Ne	
3	Na	Mg	3	4	5	6	7	8	9	10	11	12	Al	Si	P	S	Cl	Ar	
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
6	Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
7	Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Uun	Uuu	Uub		Uuq	Uuh			Uuo	
			La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu		
			Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr		

Periodic Table



# Position of elements from its electron configuration:

➤ E.g:

Classify the following elements into its appropriate group, period and block.

P :  $1s^2 2s^2 2p^6 3s^2 3p^6$

Q:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

R :  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

S :  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

T :  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

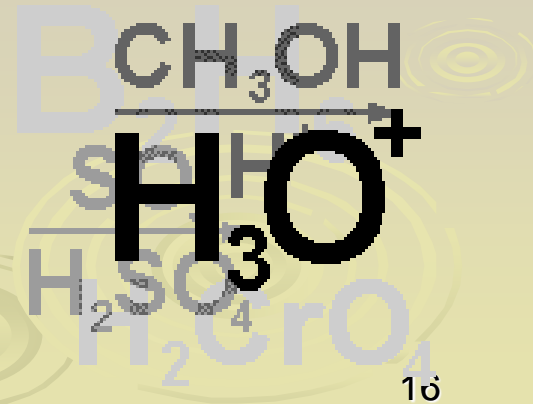
# Answer

Element	Group	Period	Class/block
P			
Q			
R			
S			
T			



# PERIODIC TABLE

## 3.2 Periodicity





# LEARNING OUTCOMES

➤ At the end of the lesson the students should be able to :

(a) Describe and analyse the variation in atomic and ionic radii

i. across a period

ii. across the first row of transition elements.

iii. down a group.

*\*Describe (C2) analyse (C3)*

(b) Compare the atomic radius of an element and its corresponding ionic radius. (C3)

(c) Define the term isoelectronic (C1)

(d) Compare the radius of isoelectronic species. (C4)

(e) Analyse the variation in the ionic radii across period 2 and 3 (C4)

## Atomic radii (nm)

	1	2	13	14	15	16	17	18
	Li 0.152	Be 0.111	B 0.088	C 0.077	N 0.070	O 0.066	F 0.064	Ne 0.070
Period 2 →	Na 0.186	Mg 0.160	Al 0.143	Si 0.117	P 0.110	S 0.104	Cl 0.099	Ar 0.094
Period 3 →	K 0.231	Ca 0.197	Ga 0.122	Ge 0.122	As 0.121	Se 0.117	Br 0.114	Kr 0.109
	Rb 0.244	Sr 0.215	In 0.162	Sn 0.140	Sb 0.141	Te 0.137	I 0.133	Xe 0.130
	Cs 0.262	Ba 0.217	Tl 0.171	Pb 0.175	Bi 0.146	Po 0.165	At 0.140	Rn 0.140

Down The group

# VARIATION IN ATOMIC AND IONIC RADII

There are two major factors affecting the size of atom in the Periodic Table:

1) Effective nuclear charge ( $Z_{\text{eff}}$ )

$$Z_{\text{eff}} = Z - S$$

$Z$  = proton number (nuclear charge)

$S$  = number of inner or core electrons

2) Shielding effect

# EFFECTIVE NUCLEAR CHARGE

- Nett charge felt by the valence electrons
- Pulls all electrons closer to nucleus
- $Z_{\text{eff}}$  increase, atomic radius decrease

# SHIELDING EFFECT

- Also known as **screening effect**.
- Caused by the **mutual repulsion** of electrons between the **inner orbital** and **valence orbital**.
- Also occur between electrons of the same orbital – but **less effective**

# VARIATION IN ATOMIC RADII

## i) Across a period

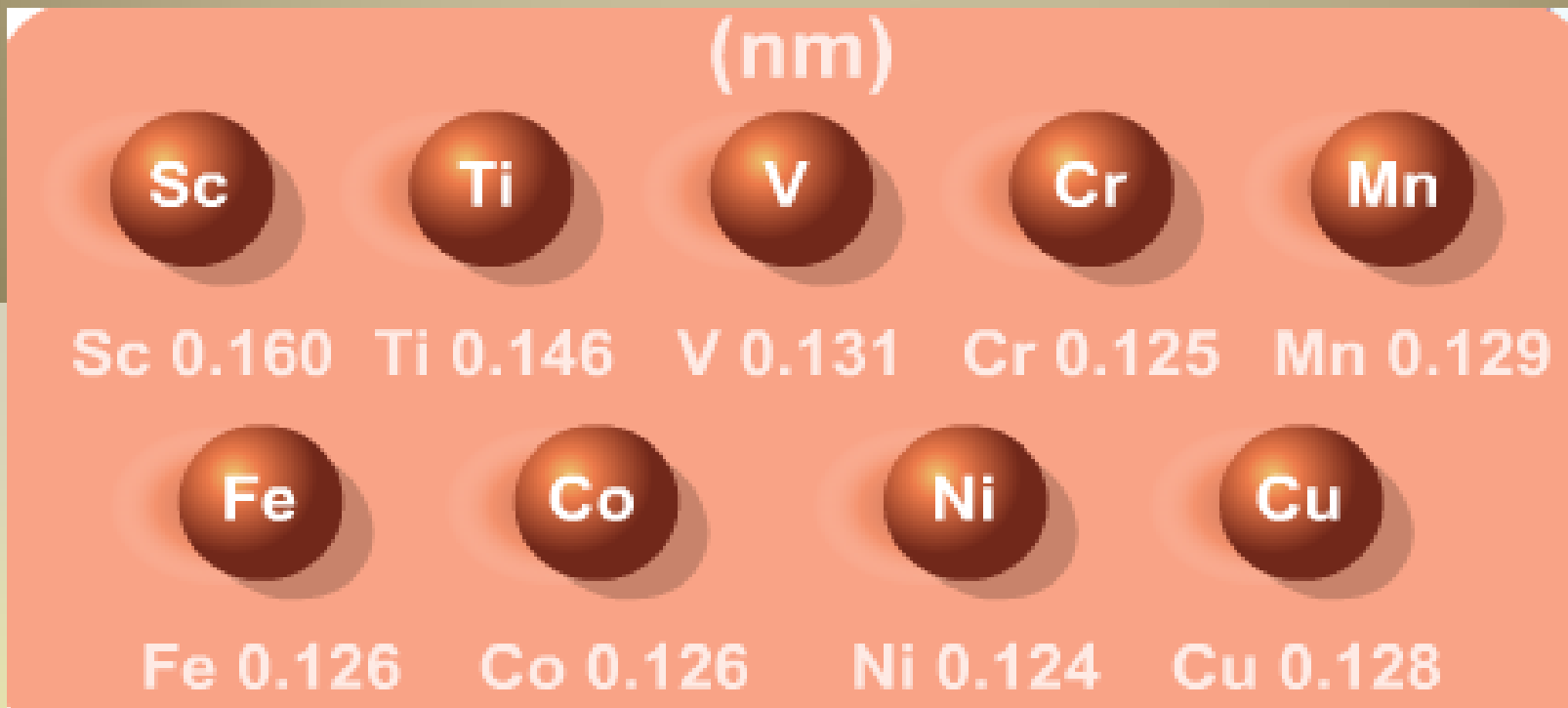
➤ The atomic radius of elements is **decreased**.

### ➤ **REASON:**

- As **proton number increases**, the **effective nuclear charge ( $Z_{\text{eff}}$ )** increases.
- Increasing the effective nuclear charge causes the **nucleus attraction towards valence electrons stronger**.
- So, valence electrons are **closer** to the nucleus
- ∴ Atomic size/radius **decreases**.

## ii) The first row of transition elements

- Atomic radius(or size) are **aproximately constant**
- Especially from Vanadium,(V) to Copper (Cu)



## REASON:

- This is because the outermost electrons are in 4s orbital.
- Across the row, electrons are added to the inner 3d orbital.
- The 3d electrons shield the outer 4s electrons from the increasing number of protons in the nucleus.
- Thus, the increase of nuclear charge is cancelled by the increase of shielding effect.
- Thus, there is no significant change in the atomic size of the transition metal.



### iii) Down a group

- Atomic radius/size **increases**.
- **REASON:**
  - Going down a group, the **no. of shell** ( $n =$  principal quantum number) **increases**.
  - No. of **inner electrons increase**.
  - So, **repulsion** between inner electrons and valence electrons **stronger**.
  - Hence, **shielding effect increases**.
  - Valence electrons are **far** from the nucleus
  - $\therefore$  atomic radius/size **increases**.

# VARIATION IN ATOMIC RADII

## NOTE:

Generally,

- **Across the period**, atomic radii decrease as  $Z_{\text{eff}}$  increases.
- **Down the group**, atomic radii increase as shielding effect increases.

# Comparison of atomic radius of an element and its corresponding ionic radius

## a) Positive ions

- Formed when an atom **loses electrons**.
- When electrons are removed from the valence shell, **electron-electron repulsions decrease**.
- **Nucleus attraction towards valence electron stronger.**
- So valence electrons are **closer** to the nucleus.
- Therefore, **cations** are always **smaller** than their corresponding atoms.

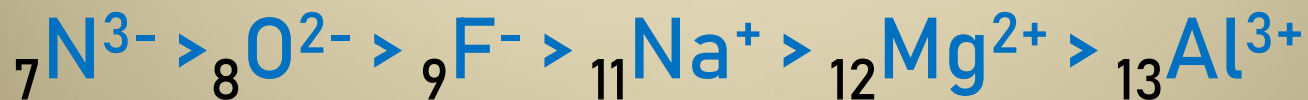
## b) Negative ions

- Formed when an atom **gains electron**.
- When electrons are gained, the **electron-electron repulsions increase**.
- Nucleus attraction towards valence electron **weaker**.
- So valence electrons are **far** from the nucleus
- Therefore, **anions** are always **larger** than the corresponding neutral atom

# The Radius of Isoelectronic Species

- Atoms or ions with the same electronic configuration are said to be isoelectronic.
- Within isoelectronic series, the **more positive** the charge, **the smaller** the species and the **more negative** the charge, **the larger** the species.
- **Reason:**  
Across a period, sizes of cations and anions **decrease** due to the **increase of  $Z_{\text{eff}}$** .

➤ Example:



Electron configuration:  $1s^2 2s^2 2p^6$

➤ Reason:

Across the period, sizes of cations and anions **decrease** due to the **increase of  $Z_{\text{eff}}$** .

# The variation in the ionic radii across period 2 and 3

## Aross Period 2:

	$\text{Li}^+$	$\text{Be}^{2+}$	$\text{B}^{3+}$		$\text{N}^{3-}$	$\text{O}^{2-}$	$\text{F}^-$
no of protons	3	4	5		7	8	9
electronic structure of atom	$1s^2 2s^1$	$1s^2 2s^2$	$1s^2 2s^2 2p^1$		$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^4$	$1s^2 2s^2 2p^5$
electronic structure of ion	$1s^2$	$1s^2$	$1s^2$		$1s^2 2s^2 2p^6$	$1s^2 2s^2 2p^6$	$1s^2 2s^2 2p^6$
ionic radius (nm)	0.060	0.031	0.020		0.171	0.140	0.133

## Across period 3:

	$\text{Na}^+$	$\text{Mg}^{2+}$	$\text{Al}^{3+}$		$\text{P}^{3-}$	$\text{S}^{2-}$	$\text{Cl}^-$
<b>no of protons</b>	11	12	13		15	16	17
<b>electronic structure of atom</b>	$1s^2 2s^2 2p^6 3s^1$	$1s^2 2s^2 2p^6 3s^2$	$1s^2 2s^2 2p^6 3s^2 3p^1$		$1s^2 2s^2 2p^6 3s^2 3p^3$	$1s^2 2s^2 2p^6 3s^2 3p^4$	$1s^2 2s^2 2p^6 3s^2 3p^5$
<b>electronic structure of ion</b>	$1s^2 2s^2 2p^6$	$1s^2 2s^2 2p^6$	$1s^2 2s^2 2p^6$		$1s^2 2s^2 2p^6 3s^2 3p^6$	$1s^2 2s^2 2p^6 3s^2 3p^6$	$1s^2 2s^2 2p^6 3s^2 3p^6$
<b>ionic radius (nm)</b>	0.102	0.072	0.054		0.212	0.184	0.181



## Positive ions across period 2:

- $\text{Li}^+ > \text{Be}^{2+} > \text{B}^{3+}$
- They are **isoelectronic species** because their **electron configurations are the same ( $1s^2$ )**

## Negative ions across period 2

- $\text{N}^{3-} > \text{O}^{2-} > \text{F}^-$
- They are **isoelectronic species with 10 electrons.**
- The electronic configuration is  $1s^2 2s^2 2p^6$

However, the proton no. increases from  $\text{Li}^+$  to  $\text{B}^{3+}$  and  $\text{N}^{3-}$  to  $\text{F}^-$ . Hence  $Z_{\text{eff}}$  increase. The electrons are pulled closer to the nucleus thus decreases the size.

## Positive ions across period 3:

- Ionic radii of  $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+}$
- They are **isoelectronic** (10 e) because their **electron configurations are the same:  $1s^2 2s^2 2p^6$ .**
- However the **no. of protons** in the nucleus of the ions is **increasing**. Hence  **$Z_{\text{eff}}$  increase**.
- That will tend to pull the electrons closer to nucleus. **Therefore the radius decrease.**

## Negative ions across period 3:

- Ionic radii of  $P^{3-} > S^{2-} > Cl^{-}$
- They are **isoelectronic (18 e)** because their electronic configurations are the same:  $1s^2 2s^2 2p^6 3s^2 3p^6$ .
- **Ionic radius decrease** because from  $P^{3-}$  to  $Cl^{-}$ , **proton no. increase**, thus  **$Z_{eff}$  increase**.

### NOTE!

There is a drastic increase in ionic radius between  $B^{3+}$  and  $N^{3-}$  and between  $Al^{3+}$  and  $P^{3-}$  due to the presence of an extra shell ( $8e^{-}$ ) for the negative ion.

# PERIODIC TABLE

## 3.2 Periodicity

	1 IA												18 8A						
1	1 H 1s <sup>1</sup>	2 He 1s <sup>2</sup>																	
2	3 Li 2s <sup>1</sup>	4 Be 2s <sup>2</sup>											5 B 2s <sup>2</sup> 2p <sup>1</sup>	6 C 2s <sup>2</sup> 2p <sup>2</sup>	7 N 2s <sup>2</sup> 2p <sup>3</sup>	8 O 2s <sup>2</sup> 2p <sup>4</sup>	9 F 2s <sup>2</sup> 2p <sup>5</sup>	10 Ne 2s <sup>2</sup> 2p <sup>6</sup>	
3	11 Na 3s <sup>1</sup>	12 Mg 3s <sup>2</sup>	3 Al 3s <sup>2</sup> 3p <sup>1</sup>	4 Si 3s <sup>2</sup> 3p <sup>2</sup>	5 P 3s <sup>2</sup> 3p <sup>3</sup>	6 S 3s <sup>2</sup> 3p <sup>4</sup>	7 Cl 3s <sup>2</sup> 3p <sup>5</sup>	8 Ar 3s <sup>2</sup> 3p <sup>6</sup>											
4	19 K 4s <sup>1</sup>	20 Ca 4s <sup>2</sup>	21 Sc 4s <sup>2</sup> 3d <sup>1</sup>	22 Ti 4s <sup>2</sup> 3d <sup>2</sup>	23 V 4s <sup>2</sup> 3d <sup>3</sup>	24 Cr 4s <sup>1</sup> 3d <sup>5</sup>	25 Mn 4s <sup>2</sup> 3d <sup>5</sup>	26 Fe 4s <sup>2</sup> 3d <sup>6</sup>	27 Co 4s <sup>2</sup> 3d <sup>7</sup>	28 Ni 4s <sup>2</sup> 3d <sup>8</sup>	29 Cu 4s <sup>1</sup> 3d <sup>10</sup>	30 Zn 4s <sup>2</sup> 3d <sup>10</sup>	31 Ga 4s <sup>2</sup> 4p <sup>1</sup>	32 Ge 4s <sup>2</sup> 4p <sup>2</sup>	33 As 4s <sup>2</sup> 4p <sup>3</sup>	34 Se 4s <sup>2</sup> 4p <sup>4</sup>	35 Br 4s <sup>2</sup> 4p <sup>5</sup>	36 Kr 4s <sup>2</sup> 4p <sup>6</sup>	
5	37 Rb 5s <sup>1</sup>	38 Sr 5s <sup>2</sup>	39 Y 5s <sup>2</sup> 4d <sup>1</sup>	40 Zr 5s <sup>2</sup> 4d <sup>2</sup>	41 Nb 5s <sup>1</sup> 4d <sup>4</sup>	42 Mo 5s <sup>1</sup> 4d <sup>5</sup>	43 Tc 5s <sup>2</sup> 4d <sup>5</sup>	44 Ru 5s <sup>1</sup> 4d <sup>7</sup>	45 Rh 5s <sup>1</sup> 4d <sup>8</sup>	46 Pd 4d <sup>10</sup>	47 Ag 5s <sup>1</sup> 4d <sup>10</sup>	48 Cd 5s <sup>2</sup> 4d <sup>10</sup>	49 In 5s <sup>2</sup> 5p <sup>1</sup>	50 Sn 5s <sup>2</sup> 5p <sup>2</sup>	51 Sb 5s <sup>2</sup> 5p <sup>3</sup>	52 Te 5s <sup>2</sup> 5p <sup>4</sup>	53 I 5s <sup>2</sup> 5p <sup>5</sup>	54 Xe 5s <sup>2</sup> 5p <sup>6</sup>	
6	55 Cs 6s <sup>1</sup>	56 Ba 6s <sup>2</sup>	57 La 6s <sup>2</sup> 5d <sup>1</sup>	72 Hf 6s <sup>2</sup> 5d <sup>2</sup>	73 Ta 6s <sup>2</sup> 5d <sup>3</sup>	74 W 6s <sup>2</sup> 5d <sup>4</sup>	75 Re 6s <sup>2</sup> 5d <sup>5</sup>	76 Os 6s <sup>2</sup> 5d <sup>6</sup>	77 Ir 6s <sup>2</sup> 5d <sup>7</sup>	78 Pt 6s <sup>1</sup> 5d <sup>9</sup>	79 Au 6s <sup>1</sup> 5d <sup>10</sup>	80 Hg 6s <sup>2</sup> 5d <sup>10</sup>	81 Tl 6s <sup>2</sup> 6p <sup>1</sup>	82 Pb 6s <sup>2</sup> 6p <sup>2</sup>	83 Bi 6s <sup>2</sup> 6p <sup>3</sup>	84 Po 6s <sup>2</sup> 6p <sup>4</sup>	85 At 6s <sup>2</sup> 6p <sup>5</sup>	86 Rn 6s <sup>2</sup> 6p <sup>6</sup>	
7	87 Fr 7s <sup>1</sup>	88 Ra 7s <sup>2</sup>	89 Ac 7s <sup>2</sup> 6d <sup>1</sup>	104 Rf 7s <sup>2</sup> 6d <sup>2</sup>	105 Db 7s <sup>2</sup> 6d <sup>3</sup>	106 Sg 7s <sup>2</sup> 6d <sup>4</sup>	107 Bh 7s <sup>2</sup> 6d <sup>5</sup>	108 Hs 7s <sup>2</sup> 6d <sup>6</sup>	109 Mt 7s <sup>2</sup> 6d <sup>7</sup>	110 Ds 7s <sup>2</sup> 6d <sup>8</sup>	111 Rg 7s <sup>2</sup> 6d <sup>9</sup>	112 Cn 7s <sup>2</sup> 6d <sup>10</sup>	(113)	114	(115)	116	(117)	118	
			58 Ce 6s <sup>2</sup> 4f <sup>5d<sup>1</sup></sup>	59 Pr 6s <sup>2</sup> 4f <sup>6</sup>	60 Nd 6s <sup>2</sup> 4f <sup>7</sup>	61 Pm 6s <sup>2</sup> 4f <sup>6</sup>	62 Sm 6s <sup>2</sup> 4f <sup>7</sup>	63 Eu 6s <sup>2</sup> 4f <sup>7</sup>	64 Gd 6s <sup>2</sup> 4f <sup>7</sup> 5d <sup>1</sup>	65 Tb 6s <sup>2</sup> 4f <sup>9</sup>	66 Dy 6s <sup>2</sup> 4f <sup>10</sup>	67 Ho 6s <sup>2</sup> 4f <sup>11</sup>	68 Er 6s <sup>2</sup> 4f <sup>12</sup>	69 Tm 6s <sup>2</sup> 4f <sup>13</sup>	70 Yb 6s <sup>2</sup> 4f <sup>14</sup>	71 Lu 6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>1</sup>			
			90 Th 7s <sup>2</sup> 6d <sup>2</sup>	91 Pa 7s <sup>2</sup> 5f <sup>6d<sup>1</sup></sup>	92 U 7s <sup>2</sup> 5f <sup>6d<sup>1</sup></sup>	93 Np 7s <sup>2</sup> 5f <sup>6d<sup>1</sup></sup>	94 Pu 7s <sup>2</sup> 5f <sup>6</sup>	95 Am 7s <sup>2</sup> 5f <sup>7</sup>	96 Cm 7s <sup>2</sup> 5f <sup>6d<sup>1</sup></sup>	97 Bk 7s <sup>2</sup> 5f <sup>9</sup>	98 Cf 7s <sup>2</sup> 5f <sup>10</sup>	99 Es 7s <sup>2</sup> 5f <sup>11</sup>	100 Fm 7s <sup>2</sup> 5f <sup>12</sup>	101 Md 7s <sup>2</sup> 5f <sup>13</sup>	102 No 7s <sup>2</sup> 5f <sup>14</sup>	103 Lr 7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>1</sup>			

# LEARNING OUTCOMES

At the end of the lesson the students should be able to :

- f) Define the first and second ionisation energies. (C1)
- g) Analyse the variations in the first ionisation energy
  - i. across a period (C4)
  - ii. down a group (C4)
- h) Explain the increase in the successive ionisation energies of an element. (C4)
- i) Deduce the **electronic configuration** of an element and its **position** in the periodic table based on successive ionisation energy data. (C4)

# IONIZATION ENERGIES (IE)

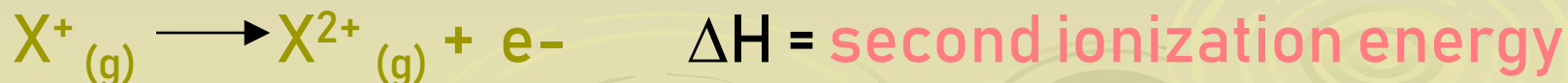
## The first ionization energy

- ⊙ The energy required to remove one mole of electron from the outermost orbital in one mole of neutral gaseous atom:



## The second ionization energy

- ⊙ The energy required to remove one mole of electron from one mole of positive ion in gaseous state:



# IONIZATION ENERGIES (IE)

- The magnitude of IE correlate with the strength of the **attractive forces between the nucleus and the outermost electron.**
- The lower of IE, the easier to remove  $e^-$ , the easier to form cation.

# FACTORS AFFECTING THE IONIZATION ENERGY

## ➤ Atomic radius

The valence electrons of an atom with a **larger radius** experience a **less attraction** towards nucleus, hence have a low ionization energy.

## ➤ Effective nuclear charge

The **higher** the nuclear **charge** the **stronger** the **attraction** between the nucleus and electrons. This causes the ionization energy to increase.

## ➤ Shielding effect ( screening effect )

The **shielding effect** of the electrons of the inner orbitals causes the outer electrons to be **less** attracted to the nucleus and thus **decrease** the magnitude of ionization energy.



# VARIATION IN THE FIRST IONIZATION ENERGY

## a) Ionization Energy Across a Period

- There is a gradual overall **increase in IE** as we move across a period, although the variation is **irregular**.
- The reason for the overall trend is the **increase in effective nuclear charge** felt by the electrons.
- This causes the electrons **closer to the nucleus** due to **nucleus attraction towards valence electron stronger**, which making it more **difficult to remove** them.

## Variation of first IE (kJ/mol) across period 2 :

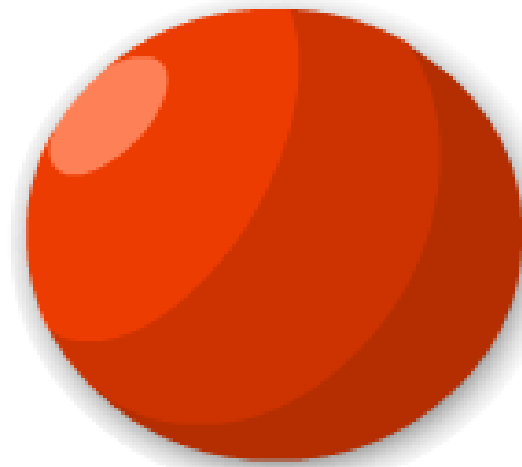
Group	1	2	13	14	15	16	17	18
Elements	Li	Be	B	C	N	O	F	Ne
IE	520	<b>900</b>	<b>801</b>	1086	<b>1402</b>	<b>1314</b>	1681	2081

- In general, **IE increases** from left to right  
but.....

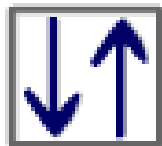
# VARIATION IN THE FIRST IONIZATION ENERGY

## Anomalous cases

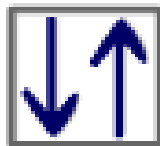
- ⊙ Between group 2 and 13. Ionization energy of **Be > B**
- ⊙ Electronic configuration of
  - Be:  $1s^2 2s^2$  (full - filled 2s orbital)
  - B:  $1s^2 2s^2 2p^1$  (partially - filled 2p orbital)
- ⊙ Full - filled orbital is more stable than partially - filled orbital
- ⊙ More energy is needed to remove the electron from the stable orbital.
- ⊙ Therefore ionization energy of **Be is higher.**



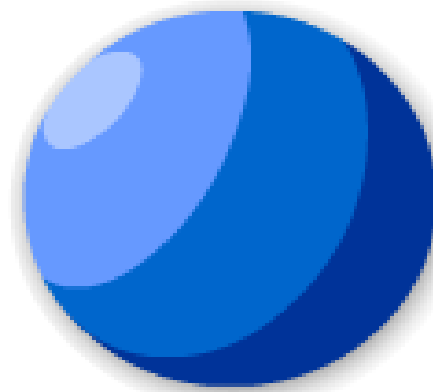
${}^4\text{Be}$



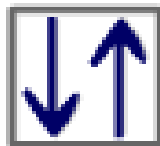
1s



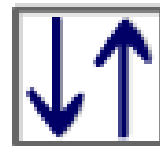
2s



${}^5\text{B}$



1s



2s



2p

Full - filled orbital

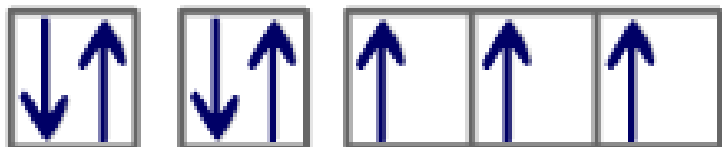
Partially - filled orbital

## Anomalous cases

- Between group 15 and 16.
- Ionization energy of **N > O**
- Electronic configuration of
  - N:  $1s^2 2s^2 2p^3$  (half - filled 2p orbital)
  - O:  $1s^2 2s^2 2p^4$  (partially - filled 2p orbital)
- Half - filled orbital is more stable than partially - filled orbital
- More energy is needed to remove the electron from the stable orbital.
- Therefore ionization energy of **N is higher.**



${}_{7}\text{N}$

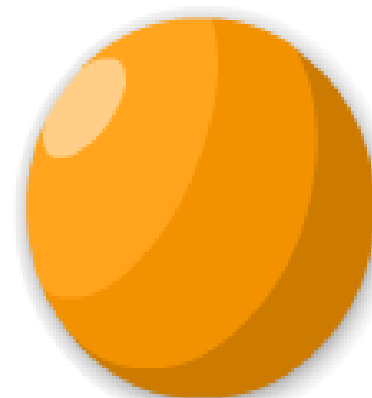


$1s^2$

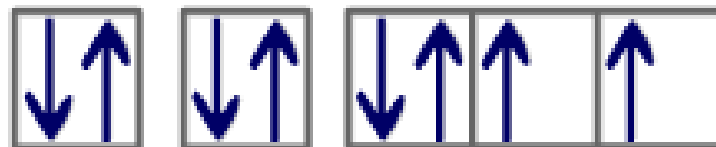
$2s^2$

$2p^3$

Half - filled orbital



${}_{8}\text{O}$



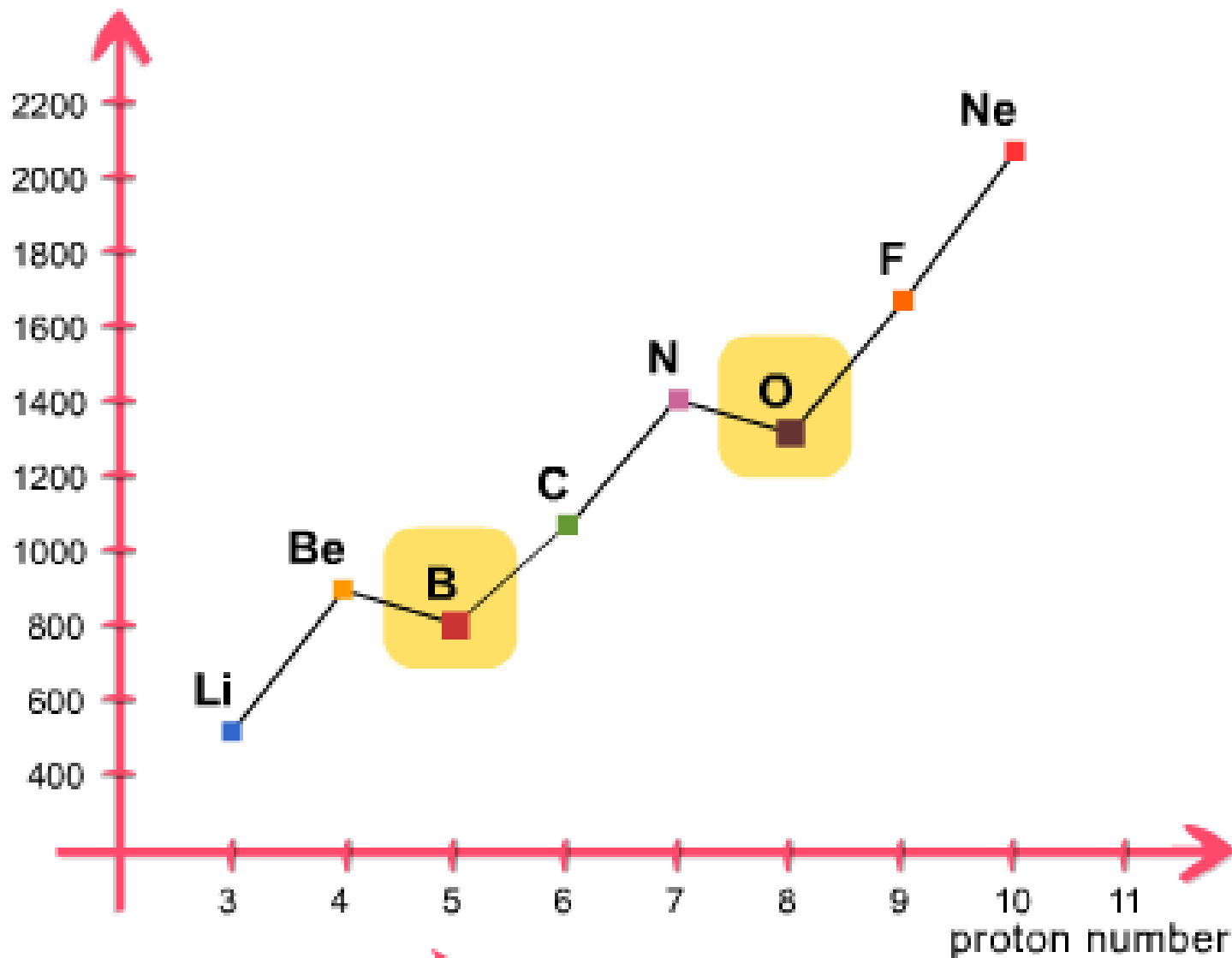
$1s^2$

$2s^2$

$2p^4$

Partially - filled orbital

Ionisation energy  $\text{kJ mol}^{-1}$



# VARIATION IN THE FIRST IONIZATION ENERGY

## b) Ionization Energy Down A Group

- Going down the group, the **shielding effect and atomic size increases**.
- the **valence electrons are far** from the nucleus **and** nucleus attraction towards valence electron weaker.
- So, **less energy** is needed to remove their first electron.
- therefore **IE decreases**.



# DETERMINATION THE VALENCE ELECTRONIC CONFIGURATION OF AN ATOM USING SUCCESSIVE IONIZATION ENERGIES

- In general, successive ionization energies always increase because each subsequent electron is being pulled away from an increasingly more positive ion, and that requires more energy.
- A large (drastic) increase in successive ionization energies occurs between the removal of the last valence electron and the removal of the first core electron (electron in the inner shell).

## Ionization energies elements (kJ mol<sup>-1</sup>)

	first	second	third	fourth	fifth	sixth	seventh
Li	520	7297	11810	-	-	-	-
Be	900	1757	14840	21000	-	-	-
B	800	2430	3659	25020	32810	-	-
C	1086	2352	4619	6221	37800	47300	-
N	1402	2857	4577	7473	9443	53250	64340
O	1314	3391	5301	7468	10980	13320	71300

- 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**.**

# IMPORTANT

- Successive ionization energy method can only determine the **group of an element** (not the period).
- Groups that involve only **1 and 2 (block s)** and **13, 14, 15, 16, 17 and 18 (block p)**.
- (transition elements are excluded)

## Example 3.2.3.

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

IE	1	2	3	4
(kJmol <sup>-1</sup> )	899	1757	14845	21000

➤ Method 1:-

By determining the IE ratios:

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

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

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

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

- 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 (8.45 times) than IE2.
- Since  $IE3 / IE2$  have the highest ratio, 2 valence electrons are present.

- Electron configuration:  $ns^2$
- This element is in Group 2



➤ Method 2:-

By determining the **differences in IE:**

$$IE_2 - IE_1 = 858 \text{ kJ mol}^{-1}$$

$$IE_3 - IE_2 = \mathbf{13088 \text{ kJ mol}^{-1}}$$

$$IE_4 - IE_3 = 6155 \text{ kJ mol}^{-1}$$



- 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 13088 kJmol<sup>-1</sup>).
- Since **IE3 - IE2 have the highest difference, 2 valence electrons are present.**
- **Electron configuration: ns<sup>2</sup>**
- **This element is Be: Group 2**



### Example 3.2.4:

Five successive ionization energies ( $\text{kJmol}^{-1}$ ) for atom M is shown below:

$\text{IE}_1$	$\text{IE}_2$	$\text{IE}_3$	$\text{IE}_4$	$\text{IE}_5$
800	1580	3230	4360	16000

Determine

- electron configuration of the valence electron for M
- group number of M in the periodic table

By determining the IE ratios:

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

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

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

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

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

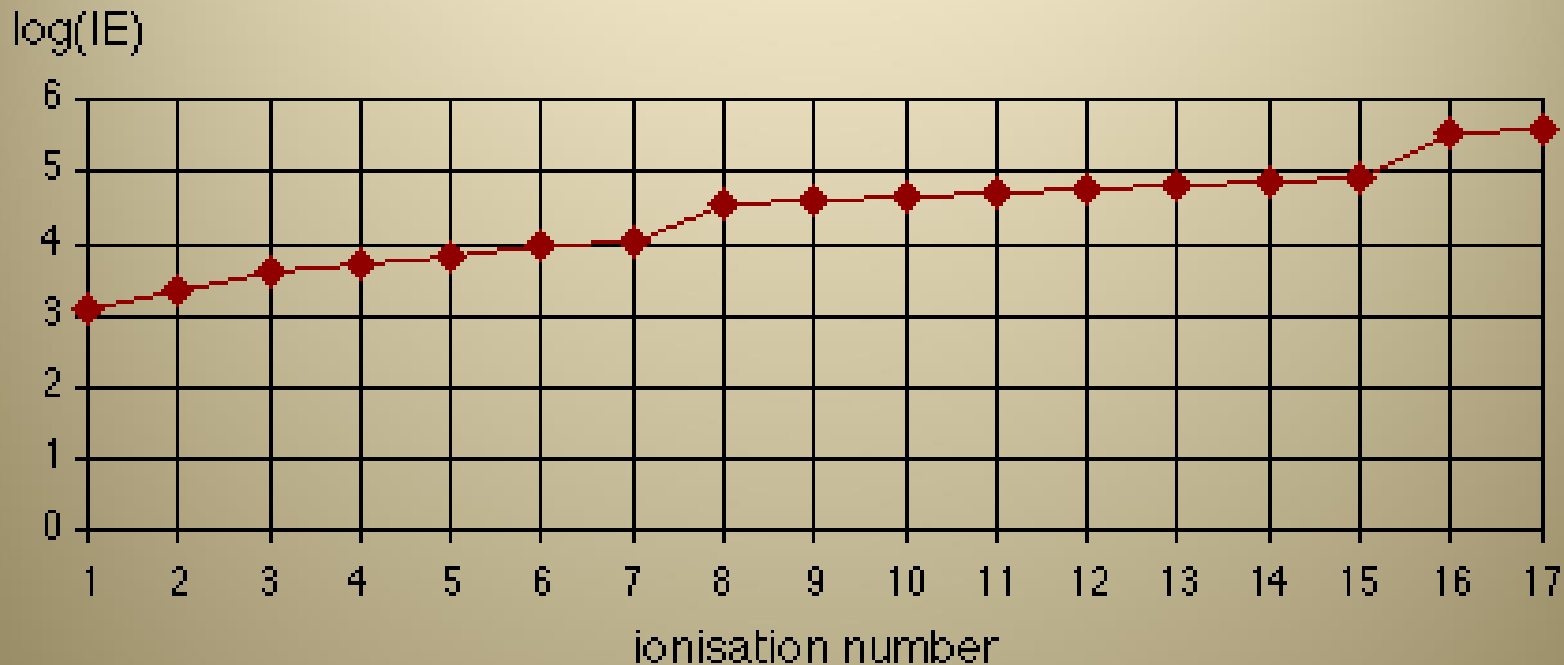
$$\frac{IE_5}{IE_4} = \frac{16000}{4360} = 3.67$$

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

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

- Since the ratio  $IE_5/IE_4$  is the highest, this shows that there are \_\_\_\_\_ in the valence shell.
- The fifth electron is removed from an inner shell with octet configuration.
- Electron configuration of valence electron for M is \_\_\_\_\_
- M is in Group \_\_\_\_ at the periodic table.

## Log graph of all the ionisation energies of chlorine



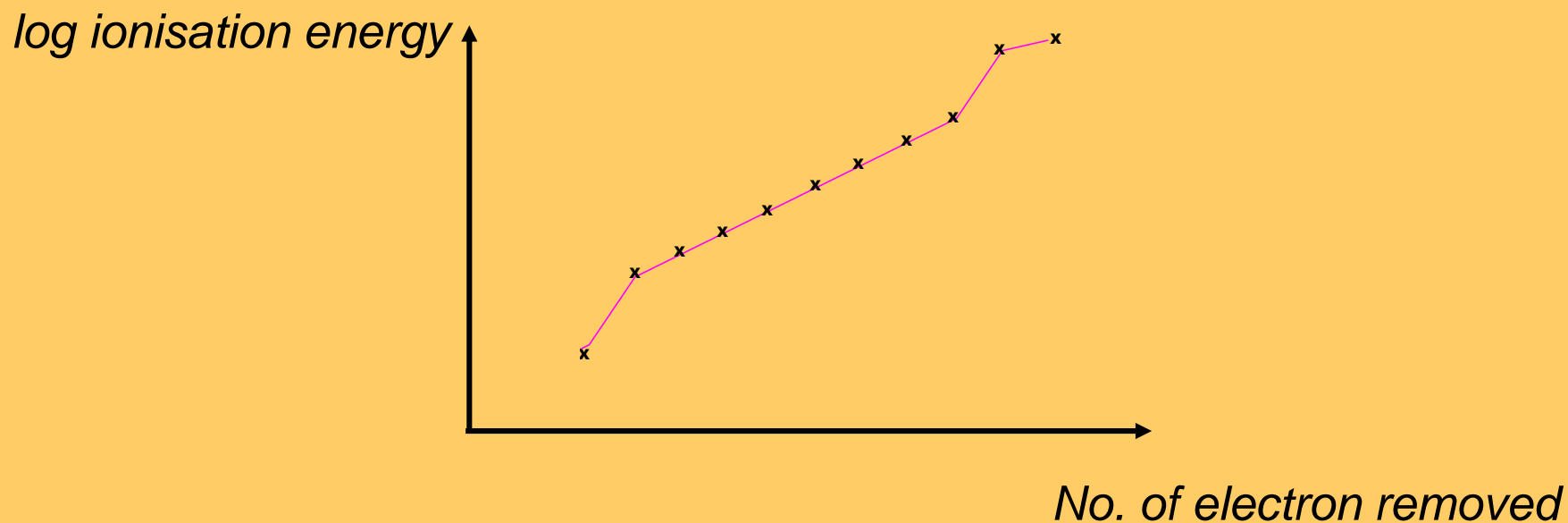
From the graph, we can determine the:

- number of electrons,
- period, and
- group

of an atom

# Exercise

The graph shows the variation of ionisation energies for element P.



State the group and period for element P.

# PERIODIC TABLE

## 3.2 Periodicity

# LEARNING OUTCOMES

At the end of the lesson the students should be able to :

- j) Define electronegativity. (C1)
- k) Explain the variation in electronegativity of elements
  - i. across a period (C2)
  - ii. down a group. (C2)
- l) Explain the acid-base character of oxides of elements in Period 3. (C3)

# ELECTRONEGATIVITY

- **Electronegativity** is the **relative tendency** of an atom **to attract electrons** to itself when chemically combined with another atom.
- Atoms with **strong attraction** for the bonding electrons have the **high electronegativity**.
- The most electronegative element is **flourine** with value 4.0.

# VARIATION IN ELECTRONEGATIVITY

## (a) Across a period

- Atomic size decreases
- Effective nuclear charge increases
- Hence, nucleus attraction towards valence electrons increases
- ∴ Electronegativity increases.



# VARIATION IN ELECTRONEGATIVITY

## (b) Down a group

- Atomic size increases
- Shielding effect increases.
- Hence, nucleus attraction towards valence electrons decreases
- ∴ Electronegativity decreases.

# VARIATION IN ELECTRONEGATIVITY

- Electronegativity *decreases down a group*, and *increases across a period*.

	1																	18
1	H 2.1	2											13	14	15	16	17	He
2	Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
3	Na 0.9	Mg 1.2	3	4	5	6	7	8	9	10	11	12	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
4	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 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr
5	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 1.9	I 2.5	Xe
6	Cs 0.7	Ba 0.9	La 1.1	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 1.8	Bi 1.9	Po 2.0	At 2.2	Rn
7	Fr 0.7	Ra 0.9	Ac 1.1															

# Exercise:

1. Arrange the elements in order of their increasing electronegativity.

a) Na, Li, Cs, K

b) B, F, Li, C

c) Cl, S, Si, Na

Answers:

a) Cs, K, Na, Li

b) Li, B, C, F

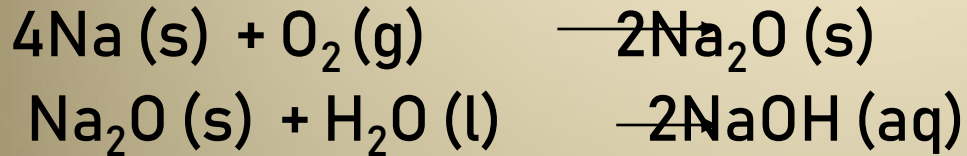
c) Na, Si, S, Cl

# ACID-BASE CHARACTER OF OXIDES OF PERIOD 3

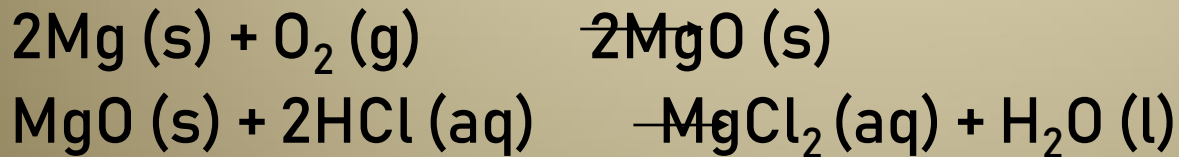
	Na <sub>2</sub> O	MgO	Al <sub>2</sub> O <sub>3</sub>	SiO <sub>2</sub>	P <sub>4</sub> O <sub>10</sub> (or P <sub>4</sub> O <sub>6</sub> )	SO <sub>3</sub> (or SO <sub>2</sub> )	Cl <sub>2</sub> O <sub>7</sub>
Adding H <sub>2</sub> O	Na <sub>2</sub> O + H <sub>2</sub> O → 2NaOH	MgO + H <sub>2</sub> O → Mg(OH) <sub>2</sub>	Insoluble	Insoluble	P <sub>4</sub> O <sub>10</sub> + 6H <sub>2</sub> O → 4H <sub>3</sub> PO <sub>4</sub>	SO <sub>3</sub> + H <sub>2</sub> O → H <sub>2</sub> SO <sub>4</sub>	Cl <sub>2</sub> O <sub>7</sub> + H <sub>2</sub> O → HClO <sub>4</sub>
Adding HCl	Na <sub>2</sub> O + H <sup>+</sup> → 2Na <sup>+</sup> + H <sub>2</sub> O	MgO + 2H <sup>+</sup> → Mg <sup>2+</sup> + H <sub>2</sub> O	Al <sub>2</sub> O <sub>3</sub> + 6H <sup>+</sup> → 2Al <sup>3+</sup> + 3H <sub>2</sub> O	No reaction	No reaction	No reaction	No reaction
Adding NaOH	No reaction	No reaction	Al <sub>2</sub> O <sub>3</sub> + 2OH <sup>-</sup> + 3H <sub>2</sub> O → 2Al(OH) <sub>4</sub>	SiO <sub>2</sub> + 2OH <sup>-</sup> → SiO <sub>3</sub> <sup>2-</sup> + H <sub>2</sub> O	P <sub>4</sub> O <sub>10</sub> + 12OH <sup>-</sup> → 4PO <sub>4</sub> <sup>3-</sup> + 6H <sub>2</sub> O	SO <sub>3</sub> + OH <sup>-</sup> → SO <sub>4</sub> <sup>2-</sup> + H <sub>2</sub> O	Cl <sub>2</sub> O <sub>7</sub> + OH <sup>-</sup> → 2ClO <sub>4</sub> <sup>-</sup> + H <sub>2</sub> O
Nature	Basic Oxide	Basic Oxide	Amphoteric Oxide	Acidic Oxide	Acidic Oxide	Acidic Oxide	Acidic Oxide

- Across a period, the oxides became more acidic.
- Basic oxides react with acid to form salt and water.
- Amphoteric oxide can react with both acid and base
- Acidic oxides react with base to form salt and water.
- Metal oxides are ionic compounds, whereas non-metallic oxides are molecular covalent compounds

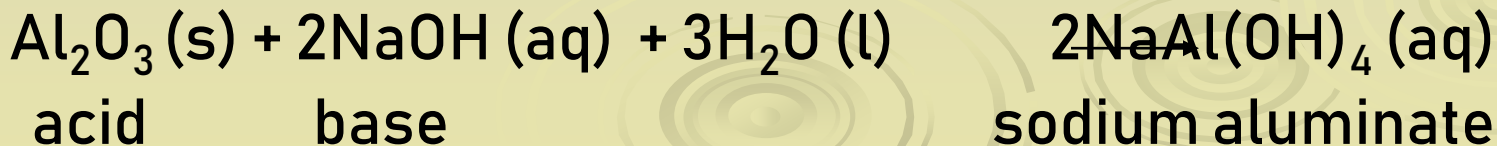
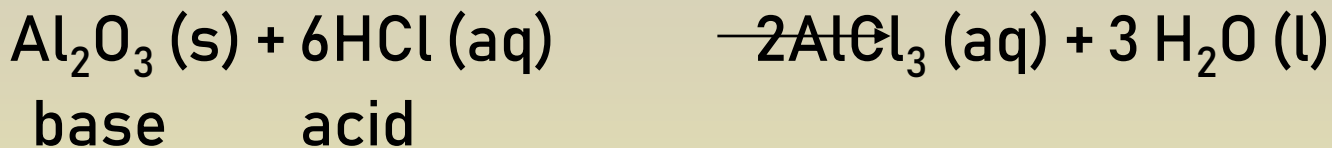
- **Na** reacts with oxygen to form **basic oxide**. The oxide will produce base when react with water.



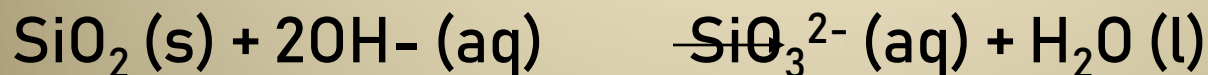
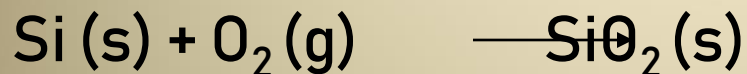
- **Mg** burns in oxygen to form **basic oxide**, MgO.



- **Al** forms **amphoteric oxide**, oxide that can react with either an acid or a base.



- **Silicon** burns in oxygen to form **acidic oxide**.

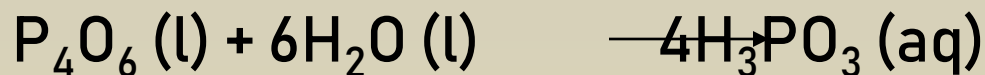


acidic

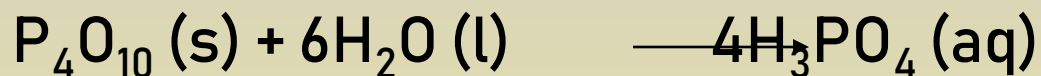
Or



- **Phosphorus** burns in oxygen to form **acidic oxide**.

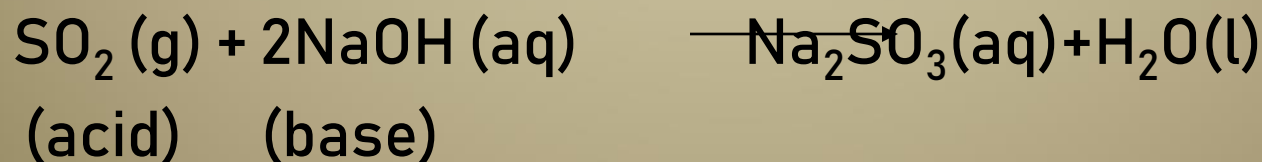
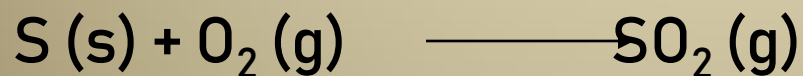


phosphorus acid



phosphoric acid

- Sulphur burns in oxygen to form acidic oxide.



- $\text{Cl}_2\text{O}_7 \text{ (g)} + \text{H}_2\text{O (l)} \longrightarrow 2\text{HClO}_4 \text{ (aq)}$   
hypochloric acid