ANDERSON SERANGOON JUNIOR COLLEGE
2025 JC2 H2 CHEMISTRY

THE PERIODIC TABLE

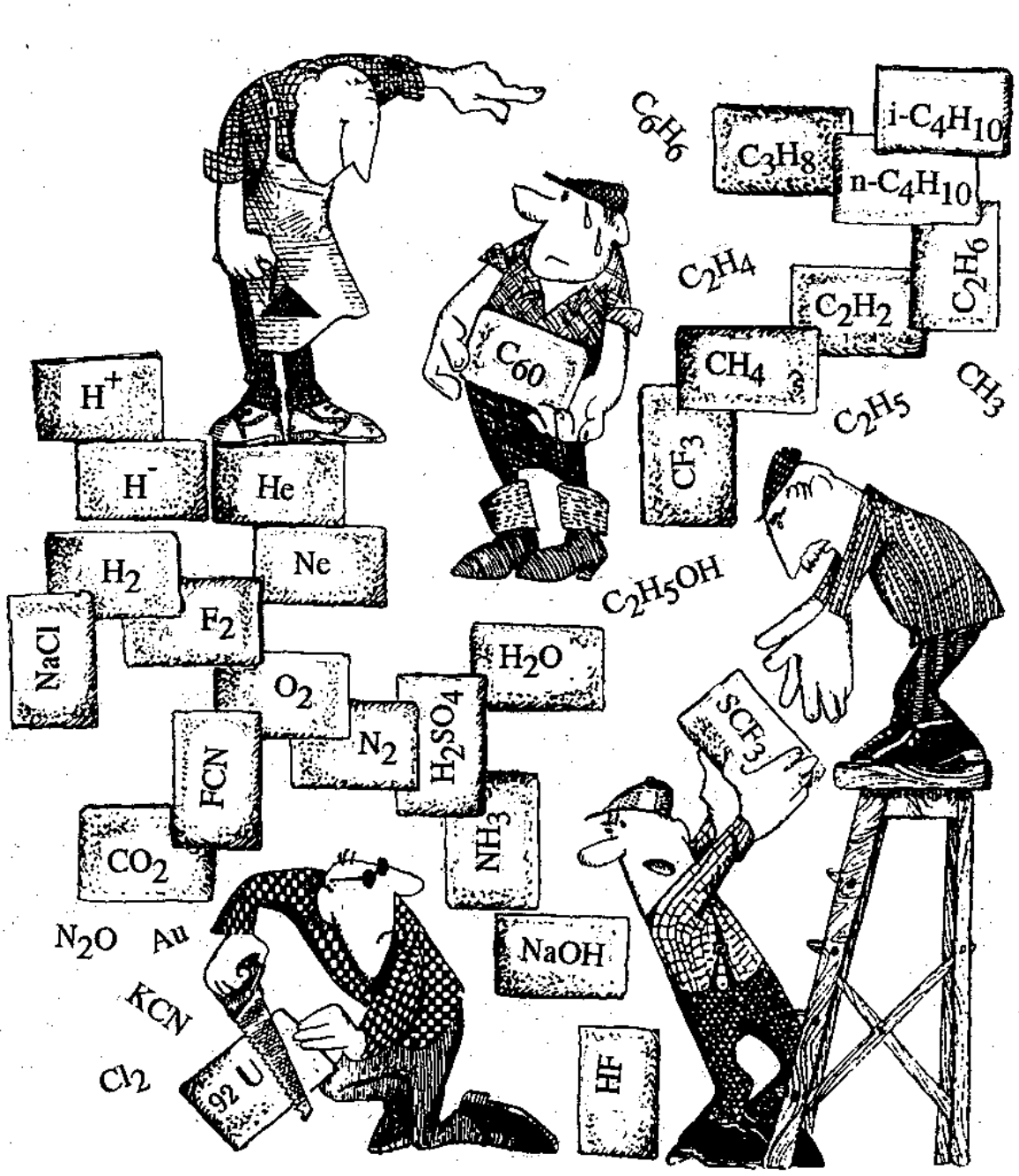

Illustration extracted from
Concept of Chemical Periodicity: from Mendeleev Table to Molecular Hyper–Periodicity Patterns
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LEARNING OBJECTIVES

Trends and variations in atomic and physical properties

For elements in the third period (sodium to chlorine), and in Group 2 (magnesium to barium) and Group 17 (chlorine to iodine):

(a) Recognize variation in the electronic configurations across a Period and down a Group

(b) Describe and explain qualitatively the trends and variations in atomic radius, ionic radius, first ionization energy and electronegativity:
  (i) Across a Period in terms of shielding and nuclear charge
  (ii) Down a Group in terms of increasing number of electronic shells and nuclear charge

(c) Interpret the variation in melting point and in electrical conductivity across a Period in terms of structure and bonding in the elements (metallic, giant molecular, or simple molecular)

(d) Describe and explain the trend in volatility of the Group 17 elements in terms of instantaneous dipole–induced dipole attraction

Trends and variations in chemical properties

For elements in the third period (sodium to chlorine):

(e) (i) State and explain the variation in the highest oxidation number of the elements in oxides (for Na_2O; MgO; Al_2O_3; SiO_2; P_4O_{10}; SO_3) and chlorides (for NaCl; $MgCl_2$; $AlCl_3$; $SiCl_4$; PCl_5)
  (ii) State and explain the variation in bonding in oxides and chlorides in terms of electronegativity (with the exception of $AlCl_3$)
  (iii) Describe the reactions of the oxides with water (for Na_2O; MgO; Al_2O_3; SiO_2; P_4O_{10}; SO_3)
  (iv) Describe and explain the acid/base behaviour of oxides (for Na_2O; MgO; Al_2O_3; SiO_2; P_4O_{10}; SO_3) and hydroxides (for NaOH; $Mg(OH)_2$; $Al(OH)_3$), including, where relevant, amphoteric behaviour in reaction with sodium hydroxide (only) and acids
  (v) Describe and explain the reactions of the chlorides with water (for NaCl; $MgCl_2$; $AlCl_3$; $SiCl_4$; PCl_5)
  (vi) Suggest the types of structure and bonding present in the oxides and chlorides from observations of their chemical and physical properties

(f) Describe and deduce from $E^{\ominus}$ values the relative reactivity of elements of:
  (i) Group 2 as reducing reagents
  (ii) Group 17 as oxidising reagents

(g) Describe and explain the trend in thermal stability of:
  (i) Group 2 carbonates in terms of the charge density of the cation and the polarisability of the large anion
  (ii) Group 17 hydrides in terms of bond energies

In addition, students should be able to:

(h) Predict the characteristic properties of an element in a given Group by using knowledge of chemical periodicity

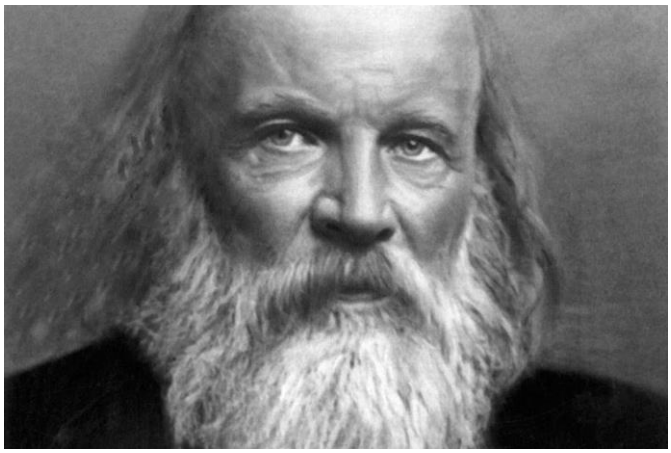
(i) Deduce the nature, possible position in the Periodic Table, and identity of unknown elements from given information of physical and chemical properties

Content Page

SDL (Watch lecture recording in MS Teams)

1. THE PERIODIC TABLE

Dmitri Mendeleev's Periodic Table, first published in 1869, which arranged elements according to **atomic mass**, was widely accepted. It was able to predict the properties of five undiscovered elements and their compounds, which provided proof of his periodic table.



The elements had been listed and carefully arranged before Mendeleev did his. They had even been organized by similar properties before.

So why is Mendeleev's periodic table the one that has endured? This was because Mendeleev correctly predicted "eka–aluminum", an element yet to be discovered in his time, until years later. All these simply from where he suspected/expected this new element was to be located on his version of the Periodic Table.

Today, there are 118 elements in the Periodic Table. Elements 113, 115, 117 and 118 newly added completing the Periodic Table. These elements were given the names nihonium (Nh), moscovium (Mc), tennessine (Ts) and oganesson (Og) respectively, officially in November 2016 by IUPAC.

A quick glance at the table shows a trend of elements in their respective groups such as electrical conductivity and reactivity. These trends, or **periodicity**, is the focus of this whole topic.

Definition
Periodicity refers to the **repeating pattern of chemical and physical properties that is seen at regular intervals**

Elements are arranged in order of **increasing atomic (proton) number**.

The properties of the elements are a periodic (i.e. **repeating after a regular interval**) function of their atomic numbers.

The Periodic Table can be segregated into **4 main blocks (s, p, d, f)**, named according to the valence orbital of the elements at ground state level (i.e. when the electrons are not excited).

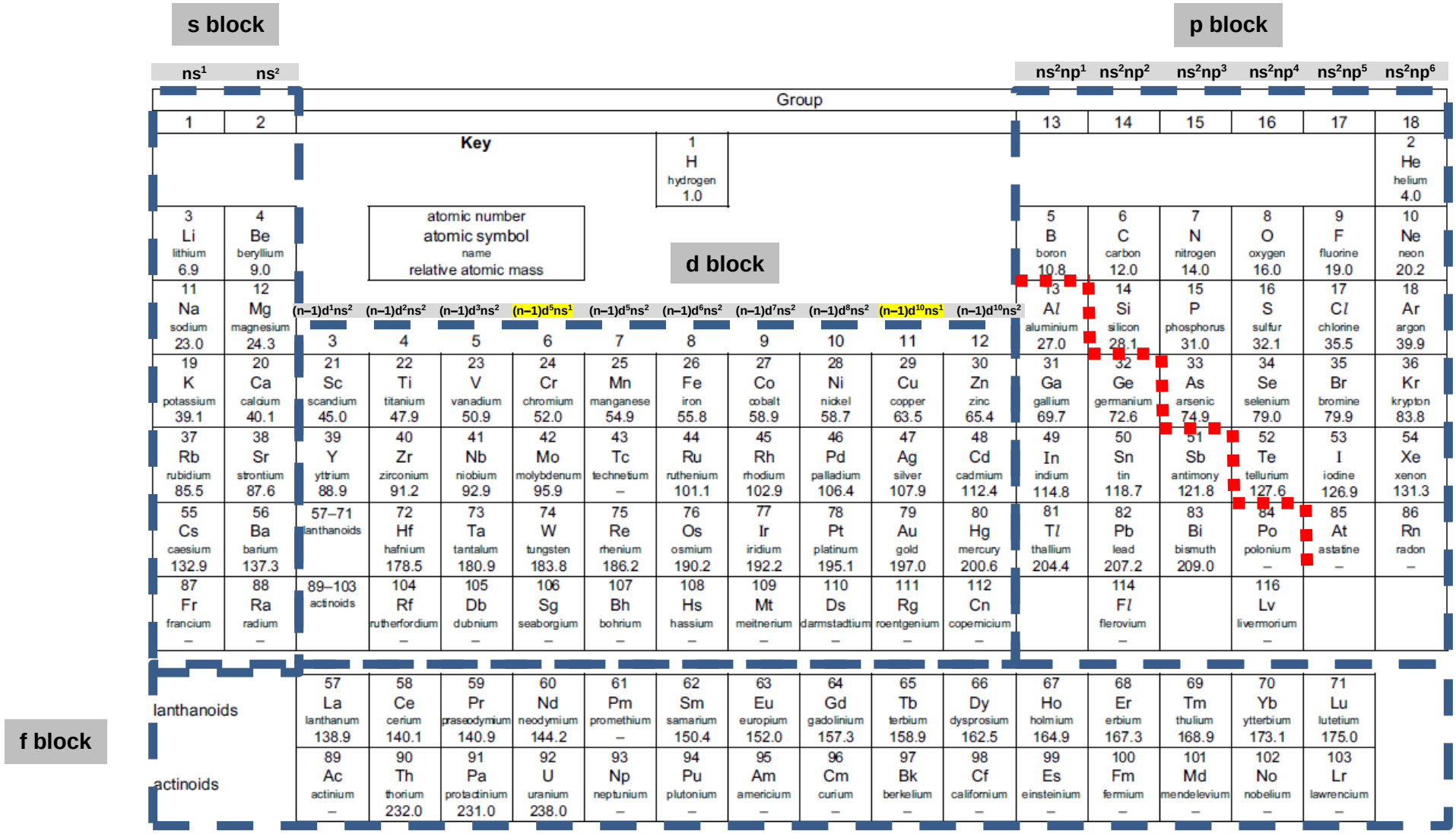
Classification of Elements in the Periodic Table:
(Refer to the Periodic Table on page 6.)

Main group elements
- Elements in Groups 1 and 2: **s–block elements**
- Elements in Groups 13 to 18: **p–block elements**

Transition elements
- Elements in Groups 3 to 12: **d–block elements**
- Elements with atomic numbers 58 – 71 & 90 – 103: **f–block elements**.

Groups refer to the vertical columns of elements in the Periodic Table.
- For Main Group elements, Group number = number of valence electrons
  (e.g. Group 13 = 3 valence electrons)
- For Transition elements, Group number = number of valence electrons in the outermost s and d subshells
  (e.g. Group 3 = 3 valence electrons in the 3d and 4s subshells)
- Elements in the same group have the **same outermost electronic configurations** in their atoms.
- Hence, they have **similar chemical properties**.

Periods refer to the horizontal rows of elements in the Periodic Table.
- Period number = principal quantum number (n) of the valence electrons
- Elements within the same period have the **same number of quantum shells.**

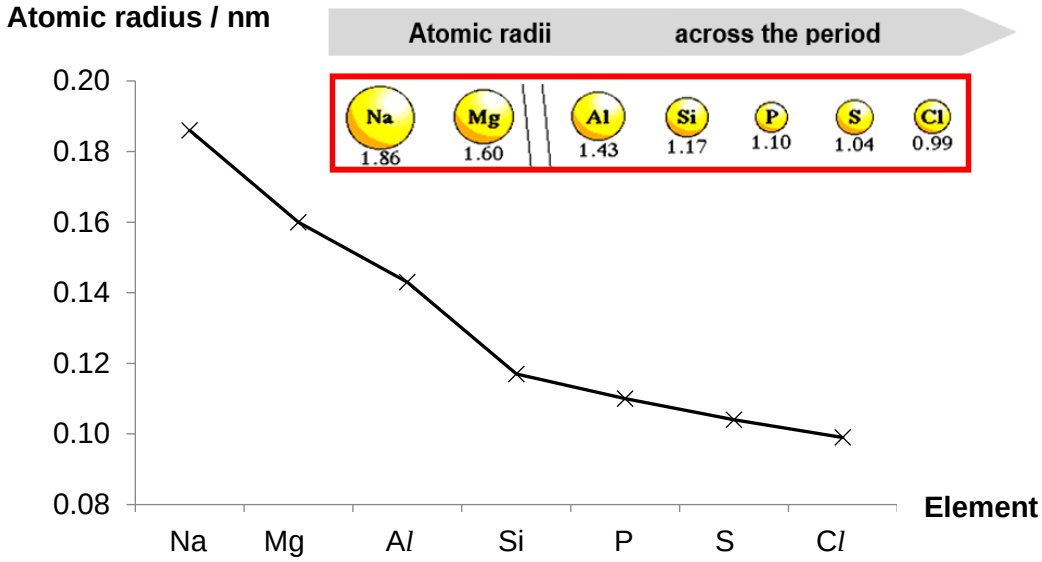

s block
p block
d block
f block
Group
Key
atomic number
atomic symbol
name
relative atomic mass
ns¹
ns²
(n–1)d¹ns²
(n–1)d²ns²
(n–1)d³ns²
(n–1)d⁵ns¹
(n–1)d⁵ns²
(n–1)d⁶ns²
(n–1)d⁷ns²
(n–1)d⁸ns²
(n–1)d¹⁰ns¹
(n–1)d¹⁰ns²
ns²np¹
ns²np²
ns²np³
ns²np⁴
ns²np⁵
ns²np⁶
1 H hydrogen 1.0
2 He helium 4.0
3 Li lithium 6.9
4 Be beryllium 9.0
5 B boron 10.8
6 C carbon 12.0
7 N nitrogen 14.0
8 O oxygen 16.0
9 F fluorine 19.0
10 Ne neon 20.2
11 Na sodium 23.0
12 Mg magnesium 24.3
13 Al aluminium 27.0
14 Si silicon 28.1
15 P phosphorus 31.0
16 S sulfur 32.1
17 Cl chlorine 35.5
18 Ar argon 39.9
19 K potassium 39.1
20 Ca calcium 40.1
21 Sc scandium 45.0
22 Ti titanium 47.9
23 V vanadium 50.9
24 Cr chromium 52.0
25 Mn manganese 54.9
26 Fe iron 55.8
27 Co cobalt 58.9
28 Ni nickel 58.7
29 Cu copper 63.5
30 Zn zinc 65.4
31 Ga gallium 69.7
32 Ge germanium 72.6
33 As arsenic 74.9
34 Se selenium 79.0
35 Br bromine 79.9
36 Kr krypton 83.8
37 Rb rubidium 85.5
38 Sr strontium 87.6
39 Y yttrium 88.9
40 Zr zirconium 91.2
41 Nb niobium 92.9
42 Mo molybdenum 95.9
43 Tc technetium –
44 Ru ruthenium 101.1
45 Rh rhodium 102.9
46 Pd palladium 106.4
47 Ag silver 107.9
48 Cd cadmium 112.4
49 In indium 114.8
50 Sn tin 118.7
51 Sb antimony 121.8
52 Te tellurium 127.6
53 I iodine 126.9
54 Xe xenon 131.3
55 Cs caesium 132.9
56 Ba barium 137.3
57–71 lanthanoids
72 Hf hafnium 178.5
73 Ta tantalum 180.9
74 W tungsten 183.8
75 Re rhenium 186.2
76 Os osmium 190.2
77 Ir iridium 192.2
78 Pt platinum 195.1
79 Au gold 197.0
80 Hg mercury 200.6
81 Tl thallium 204.4
82 Pb lead 207.2
83 Bi bismuth 209.0
84 Po polonium –
85 At astatine –
86 Rn radon –
87 Fr francium –
88 Ra radium –
89–103 actinoids
104 Rf rutherfordium –
105 Db dubnium –
106 Sg seaborgium –
107 Bh bohrium –
108 Hs hassium –
109 Mt meitnerium –
110 Ds darmstadtium –
111 Rg roentgenium –
112 Cn copernicium –
114 Fl flerovium –
116 Lv livermorium –
lanthanoids
57 La lanthanum 138.9
58 Ce cerium 140.1
59 Pr praseodymium 140.9
60 Nd neodymium 144.2
61 Pm promethium –
62 Sm samarium 150.4
63 Eu europium 152.0
64 Gd gadolinium 157.3
65 Tb terbium 158.9
66 Dy dysprosium 162.5
67 Ho holmium 164.9
68 Er erbium 167.3
69 Tm thulium 168.9
70 Yb ytterbium 173.1
71 Lu lutetium 175.0
actinoids
89 Ac actinium –
90 Th thorium 232.0
91 Pa protactinium 231.0
92 U uranium 238.0
93 Np neptunium –
94 Pu plutonium –
95 Am americium –
96 Cm curium –
97 Bk berkelium –
98 Cf californium –
99 Es einsteinium –
100 Fm fermium –
101 Md mendelevium –
102 No nobelium –
103 Lr lawrencium –

2. TRENDS OF PHYSICAL PROPERTIES IN THE PERIODIC TABLE *(SDL)*

Recall the topic on 'Atomic Structure' how **nuclear charge, shielding effect and number of electronic shells of an atom** affect its ionisation energy. In this section, you will further learn how these three factors affect some other physical properties of an element.

2.1 Atomic Radius *(SDL)*

Atomic radius is defined as the **half the distance** between the **nuclei of two adjacent atoms** of the same element.



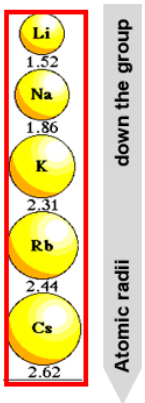
Note: Values of atomic radius are given in angstrom (1×10^{-10} m)

Across the period,

- **Nuclear charge increases**.
- Number of electrons also increases but these electrons are added to the same outermost electron shell and hence **number of electron shells remain the same** and **shielding effect remains approximately constant.**
- There are **stronger** (electrostatic) **forces of attraction** between the nucleus and the valence electrons.
- **Atomic radius decreases** across the period.

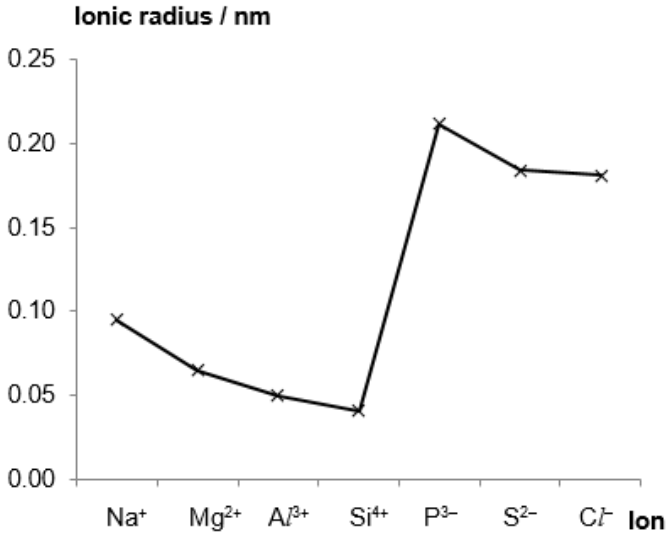
Down the group,

- **Nuclear charge increases.**
- The number of filled principal quantum shells **increases**.
- The valence electrons are **further** from the nucleus and **more shielded**.
- There are **weaker** (electrostatic) **forces of attraction** between the nucleus and the valence electrons.
- **Atomic radius increases** down the group.



2.2 Ionic Radius *(SDL)*

Ionic radius is defined as the **half the distance** between the **nuclei of two adjacent ions** of the same element.



From Na^+ to Si^{4+} and from P^{3-} to Cl^-

- **Nuclear charge increases**.
- **Number of filled principal quantum shells and shielding effect remains the same** since the cations (Na^+ to Si^{4+}) and anions (P^{3-} to Cl^-) are **isoelectronic**.
- There are **stronger** (electrostatic) **forces of attraction** between the nucleus and the valence electrons.
- **Ionic radius decreases**.

From Si^{4+} to P^{3-},

Si^{4+} : $1s^2\ 2s^2\ 2p^6$ $\quad$ P^{3-} : $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^2\ 3p^6}$

- **Nuclear charge increases**
- P^{3-} has **one more filled principal quantum shell** compared to Si^{4+}.
- The valence electrons in P^{3-} are **further** from the nucleus than in Si^{4+} and **more shielded.**
- There are **weaker** (electrostatic) **forces of attraction** between the nucleus and the valence electrons in P^{3-} than in Si^{4+}.
- Hence ionic radius of P^{3-} is greater than Si^{4+}.

Comparison of Atomic Radius and Ionic Radius

From the Data Booklet,

	Na	**Mg**	**Al**	**Si**
Proton No	11	12	13	14
Electronic Config	[Ne] $3s^1$	[Ne] $3s^2$	[Ne] $3s^23p^1$	[Ne] $3s^23p^2$
atomic radius /nm	0.186	0.160	0.143	0.117
	Na^+	**Mg^{2+}**	**Al^{3+}**	**Si^{4+}**
Proton No	11	12	13	14
Electronic Config	[He] $2s^22p^6$	[He] $2s^22p^6$	[He] $2s^22p^6$	[He] $2s^22p^6$
cationic radius /nm	0.095	0.065	0.050	0.041

Cationic radius is <u>smaller</u> *than* atomic radius.

- **Nuclear charge** remains the **same**.
- The cation has **one principal quantum shell lesser** than the atom.
- The valence electrons in the cation are **closer** to the nucleus and **less shielded** from the nucleus.
- There are **stronger** (electrostatic) **forces of attraction** between the nucleus and the valence electrons in cation.

(2)	**P**	**S**	**Cl**
Proton No	15	16	17
Electronic Config	[Ne] $3s^23p^3$	[Ne] $3s^23p^4$	[Ne] $3s^23p^5$
atomic radius /nm	0.110	0.104	0.099
	P^{3-}	**S^{2-}**	**Cl^-**
Proton No	15	16	17
Electronic Config	[Ne] $3s^23p^6$	[Ne] $3s^23p^6$	[Ne] $3s^23p^6$
anionic radius /nm	0.212	0.184	0.181

Anionic radius is <u>larger</u> than atomic radius.

- **Nuclear charge** remains the **same**.
- The anion has **more electrons** than the atom.
- **Inter–electronic repulsion increases**, resulting in the valence electrons in being **further** away from the nucleus.
- There are **weaker** (electrostatic) **forces of attraction** between the nucleus and the valence electrons in anion.

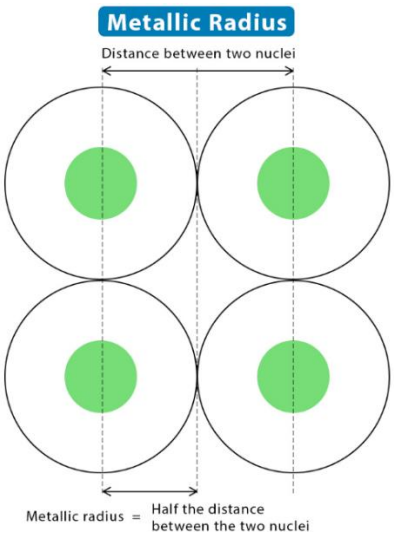
Good to know:

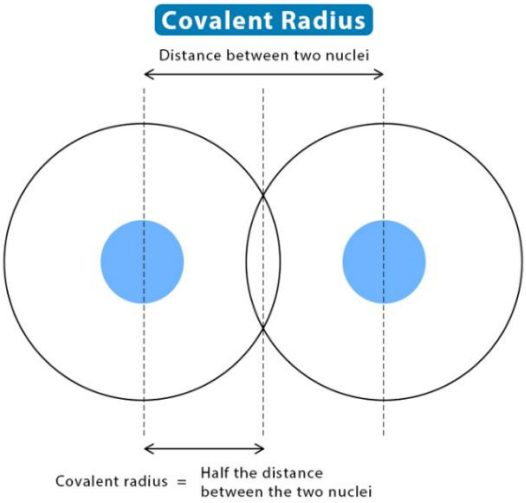
Let's find out why is the radius of Ar left out in the comparison for atomic radius?

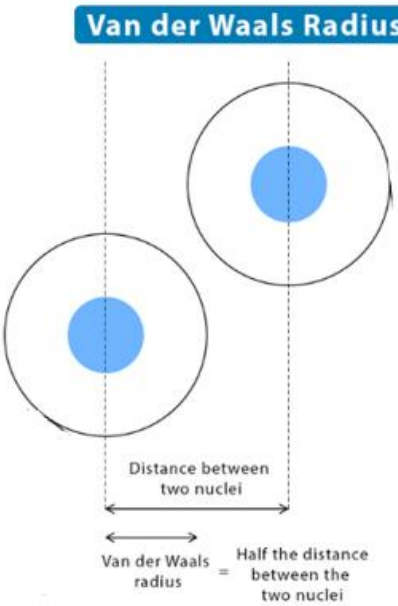
There are three different measurements for atomic radius.

1. Metallic radius
2. Covalent radius
3. Van der Waals' radius

These measurements can be found in *Table 5 of the Data Booklet*. Radii of atoms can be different using different measurements. Covalent radius of the atom is smaller than its Van der Waals' radius for the same atom.







For noble gases, we tend to refer their radius as the van der Waals' radius. Noble gases exist in monoatomic form and we refer to their radius as the distance between the nuclei of 2 atoms. The proximity of the atoms **is not as great as in metals** as there is a large space in between the atoms, due to the presence of weak Van der Waals forces that attracts the atoms together.

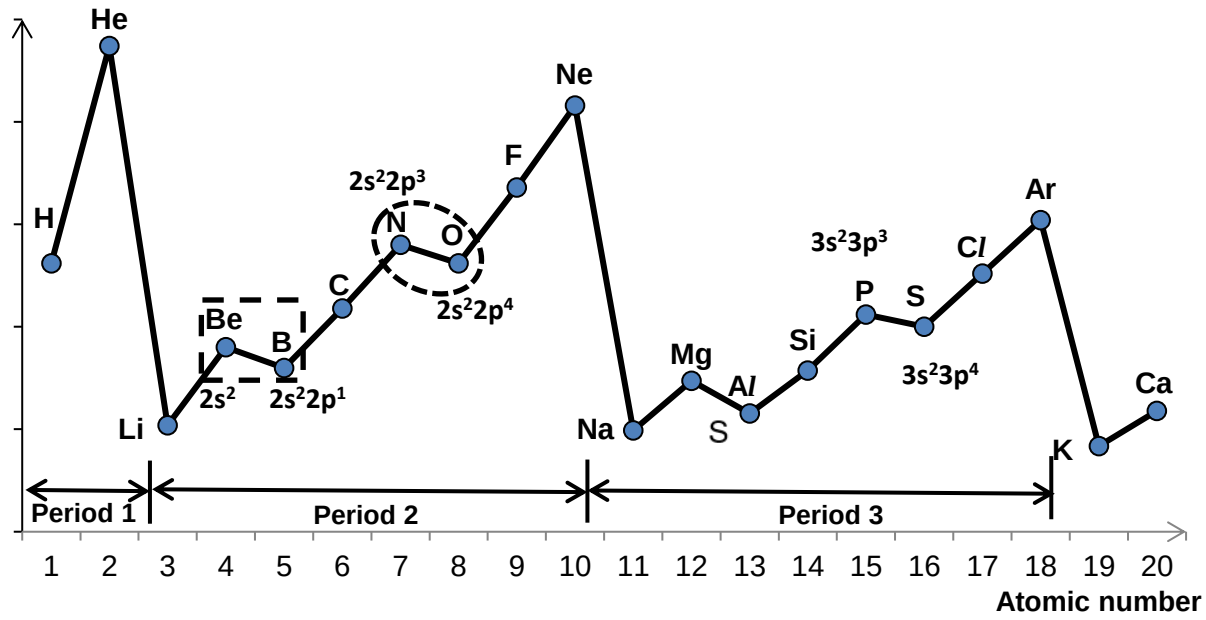
(Note: Van der Waals forces are similar to instantaneous dipole–induced dipole attractions. However, do not use the term 'Van der Waals forces' in examination.)

2.3 1st Ionisation Energy *(Recall from 'The Atomic Structure') (SDL)*

First ionisation Energy (1st IE) is the energy required to remove **one mole of electrons** from **one mole of gaseous atoms** of the element to form **one mole of singly positively charged gaseous ions**.

$$X\,(g) \rightarrow X^{+}\,(g) + e^{-} \qquad \Delta H > 0$$

First I.E. / kJ mol^{-1}



(1) General trend across the period from Na to Ar,
• **Nuclear charge increases**. • Number of electrons also increases but these electrons are added to the same outermost electron shell and hence **number of electron shell remain the same** and **shielding effect remains approximately constant.** • **More energy** is required to overcome the **stronger** (electrostatic) **forces of attraction** between the nucleus and the valence electron to be removed. • First ionisation energies **generally increase** across a period.
(2) Anomalies across the period
The first I.E. of Al (577 kJ mol^{-}) is **lower** than that of Mg (736 kJ mol^{-1}). Mg: $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^2}$ Al: $1s^2\ 2s^2\ 2p^6\ 3s^2\ \mathbf{3p^1}$ • The **3p electron** to be removed from Al is at a **higher energy level** and **less strongly attracted to the nucleus** than the **3s electron** to be removed from Mg. • Hence **less energy** is required to remove the **3p electron** in Al than the **3s electron** in Mg. • Consequently the first ionisation energy of Al is **lower** than that of Mg.

(3) Anomalies across the period

The first I.E of S (1000 kJ mol^{-1}) is **lower** than that of P (1060 kJ mol^{-1}).

P : $1s^2\ 2s^2\ 2p^6\ 3s^2$ **$3p^3$** S : $1s^2\ 2s^2\ 2p^6\ 3s^2$ **$3p^4$**

- The **3p electron** to be removed from S is a **paired electron** while that to be removed from P is an **unpaired electron**.
- Due to **inter–electronic repulsion** between paired electrons in the same orbital, **less energy** is required to remove the 3p electron from S.
- Hence the first ionisation energy of S is **lower** than that of P.

(4) Down the group

- **Nuclear charge increases**
- The number of filled principal quantum shells **increases**.
- The valence electrons are **further** from the nucleus and **more shielded**.
- **Less energy** is required to overcome the **weaker** (electrostatic) **forces of attraction** between the nucleus and the valence electron to be removed.
- First ionisation energies **decrease** down the group.

2.4 Electronegativity *(Recall 'Chemical Bonding' Page 14) (SDL)*

Electronegativity refers to the ability of an atom to **attract the shared pair(s) of electrons** in a **covalent bond**.

Electronegativity INCREASES across the period

Electronegativity DECREASES down the group

Group (vertical)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Period (horizontal)																		
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo
Lanthanides	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
Actinides	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.291		

Periodic table of electronegativity using the Pauling scale

(1) Across the period

- **Nuclear charge increases**.
- Across a period, number of electrons also increases but these electrons are added to the same outermost electron shell and hence **number of electron shells remains the same** and **shielding effect remains approximately constant**.
- There are **stronger** (electrostatic) **forces of attraction** between the **nucleus and the electron pair in a covalent bond.**
- Electronegativity **increases.**

(2) Down the group

- **Nuclear charge increases**
- The number of filled principal quantum shells **increases**.
- The valence electrons are **further** from the nucleus and **more shielded**.
- There are **weaker** (electrostatic) **forces of attraction** between the **nucleus and the electron pair in a covalent bond.**
- Electronegativity **decreases.**

(3) Diagonal Relationship (IMPORTANT)

- The **head element (Li, Be, B)** of each group may **resemble** the elements positioned **diagonally** to it, **both physically and chemically.**

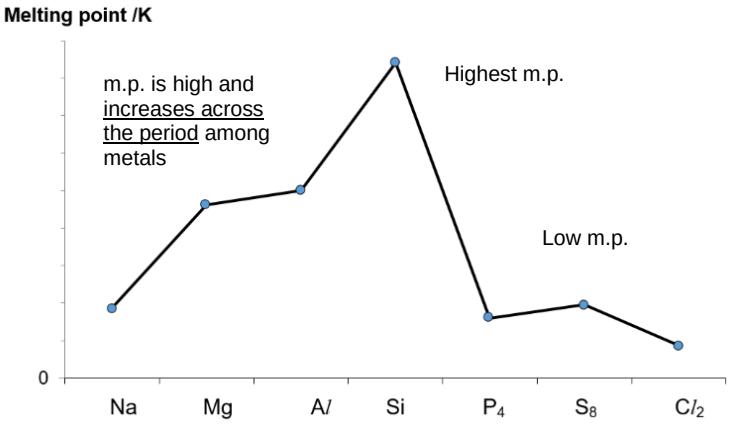
- Due to *increase in electronegativity across the period but decrease down the group*, the elements diagonal to each other can have **similar electronegativity** giving rise to similar properties.

Group	1	2	13	14
Period 2	Li (0.98) (i)	Be (1.57) (ii)	B (2.04) (iii)	C (2.55)
Period 3	Na (0.93)	Mg (1.31)	A*l* (1.61)	Si (1.90)

Note: Values in brackets are the electronegativity values of the elements.

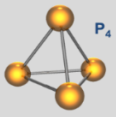
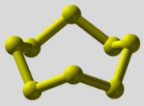
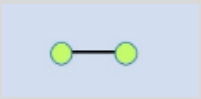
	Some similar properties that are exhibited by elements diagonally in the Periodic Table
(i)	Unlike all other carbonates of Group 1 elements (e.g. Na_2CO_3, K_2CO_3), **Li_2CO_3** undergoes similar **thermal decomposition** as **$MgCO_3$**. Li_2CO_3 (s) → Li_2O (s) + CO_2 (g) $\quad\quad$ $MgCO_3$ (s) → MgO (s) + CO_2 (g) The rest of the Group 1 carbonates (Na_2CO_3, K_2CO_3) are thermally stable.
(ii)	Unlike all other Group 2 elements, beryllium hydroxides exhibits **amphoteric character** and can **react with acids and alkalis** (e.g. sodium hydroxide) in the same way as **aluminium hydroxides.** • Both $Be(OH)_2$ and A*l*$(OH)_3$ are **amphoteric**. *With acid* $Be(OH)_2 + 2H^+ \rightarrow Be^{2+} + 2H_2O$ $\quad\quad$ A*l*$(OH)_3 + 3H^+ \rightarrow$ A*l*$^{3+} + 3H_2O$ *With alkali* $Be(OH)_2 + OH^- \rightarrow Be(OH)_3^-$ $\quad\quad$ A*l*$(OH)_3 + OH^- \rightarrow$ A*l*$(OH)_4^-$ • Be^{2+}(aq) and A*l*$^{3+}$(aq) are **weakly acidic** due to **hydrolysis**. (The charge density $\frac{q^+}{r^+}$ of Be^{2+}(aq) and A*l*$^{3+}$(aq) is relatively high.) $[Be(H_2O)_4]^{2+}$ (aq) ⇌ $[Be(H_2O)_3(OH)]^+$ (aq) + H^+ (aq) $[$A*l*$(H_2O)_6]^{3+}$ (aq) ⇌ $[$A*l*$(H_2O)_5(OH)]^{2+}$ (aq) + H^+ (aq)
(iii)	Both B and Si are used as **semiconductors**.

2.5 Melting and Boiling Points of Period 3 Elements *(Recall 'Chemical Bonding') (SDL)*



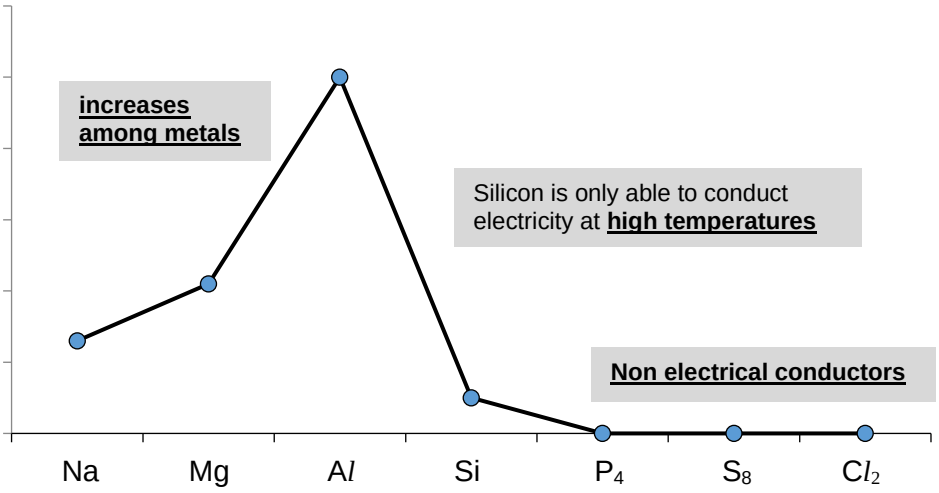
Elements	**Na** (97.79 °C)	**Mg** (650 °C)	**Al** (660 °C)
Melting point	Metals have **high melting points** and **increases** among the metals		
Explanation	Metals have **giant metallic structure.** ○ **Number of valence electrons** per atom for metallic bonding **increases**, and **cationic radii decreases** ○ **Metallic bond strength increases** ○ **Large amount of energy** is required to overcome the **stronger** (electrostatic) forces of attraction between **cations** and **sea of delocalised electrons.** ○ Melting point is **high** and **increases** across the period among the metals.		

Elements	**Si** (1414°C)
Melting point	**Very high**
Explanation	Si has a **giant molecular structure** **Very large amount of energy** required to overcome the **extensive network of strong covalent bonds** between atoms in the **three dimensional network structure**.

Elements	**P_4** (111.57 °C)	**S_8** (115.2 °C)	**Cl_2** (–105.5 °C)
			
Melting point	○ Melting point is generally **low**. ○ **Melting point: S_8 > P_4 > Cl_2** ○ *Trend follows the decreasing number of electrons*		
Explanation			
Useful for comparing elements with different structures like metals and simple molecules.	P_4, S_8 and Cl_2 are **simple molecules** with **low melting points** **Small amount of energy** is required to overcome the **weak instantaneous dipole–induced dipole attractions** between the molecules.		
Useful for comparing elements with same structures.	**Simple molecules** with melting point follows the **trend S_8 > P_4 > Cl_2** **Number of electrons** of S_8 > P_4 > Cl_2 **Amount of energy** required to **overcome decreasing strength of instantaneous dipole–induced dipole attractions decreases** between the molecules from S_8 > P_4 > Cl_2		

2.6 Electrical Conductivity of Period 3 Elements (SDL)

Electrical Conductivity



Element	Na	Mg	Al
Electrical conductivity	**Good conductors of electricity** Electrical conductivity **increases** across the period among the metals		
Explanation	**Giant metallic structure** Presence of **sea of delocalised electrons** as **mobile charge carriers** to conduct electricity under the influence of an electric current/ field. **Electrical conductivity increases across the period among the metals** **Number of valence electrons** contributed per atom for metallic bonding **increases** across the period among the metals for **delocalisation**. **More mobile electrons** can act as **charge carriers.**		

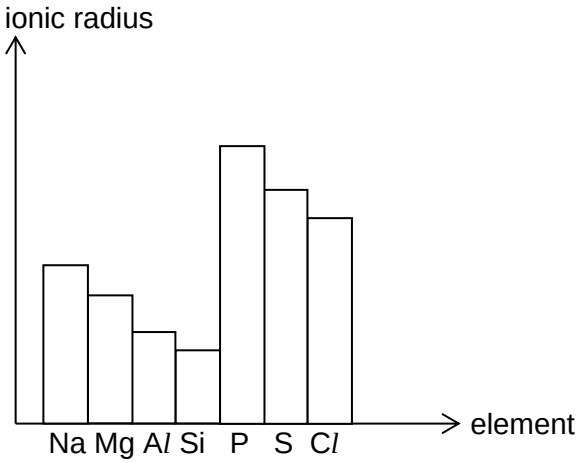
Elements	Si
Electrical conductivity	Si is a **semiconductor** which can only give rise to delocalised electrons at **high temperatures**
Explanation	**Giant molecular structure** Strong covalent bonds exist **between atoms** in a giant molecular structure At **higher** temperatures, more electrons **gain enough energy** to **overcome the nuclear attraction** and electrical conductivity **increases**

Elements	P_4	S_8	Cl_2
Electrical conductivity	**Non–conductors of electricity**		
Explanation	**Simple molecular structure** **Absence** of delocalised electrons or free mobile ions as **mobile charge carriers** in their molecular structures**.**		

Checkpoint 1

5

1. The graph below shows the ionic radii of seven elements found in Period 3.



Which of the following statements correctly explains the trends shown in the graph?

A The ionic radius of P^{3-} ion is greater than that of Si^{4+} ion due to P^{3-} having more protons than Si^{4+}.

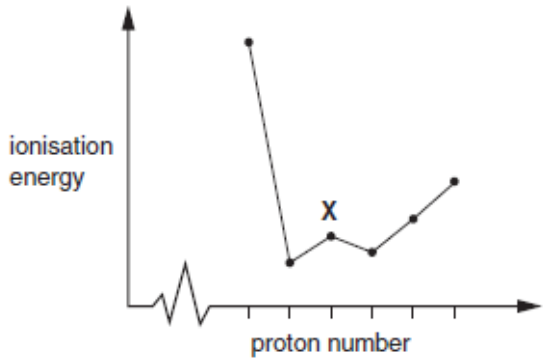
B The ionic radius decreases from Na^{+} ion to Si^{4+} ion due to decreasing shielding effect by inner shell electrons.

C The ionic radius decreases from Na^{+} ion to Si^{4+} ion due to decreasing nuclear charge.

D The ionic radius decreases from P^{3-} ion to Cl^{-} ion due to increasing nuclear charge.

( )

2. The sketch below shows the variation of first ionisation energy with proton number for six elements of consecutive proton numbers between 1 and 18 (H to Ar).



What is the identity of element **X**?

A Mg **B** Al **C** Si **D** P

( )

3. Which of the following statements are correct for the sequence of compounds below considered from left to right?

NaF MgO A*l*N SiC

A Electronegativity difference between the elements in each compound decreases.

B The formula–units of these compounds are not iso–electronic (having the same number of electrons).

C The bonding becomes decreasingly covalent due to the increasing electronegativity differences.

D Electrical conductivity increases when dissolved in solution.

( )

4. Group 2 metals have lower melting points than the Group 13 metals. Which of the following factors can explain the above phenomena?

A The atomic radius of Group 2 metals is smaller.

B There are two valence electrons delocalised into the metallic lattice of Group 2 metals.

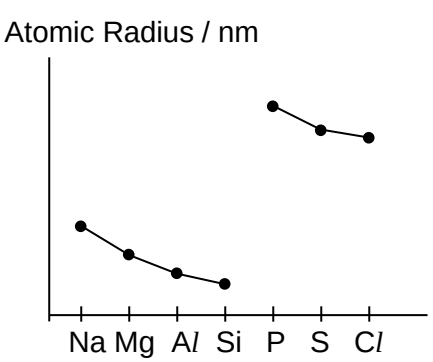
C Group 2 metals have lower first ionisation energies.

D The interatomic distances in the metal lattices of Group 2 metals are smaller.

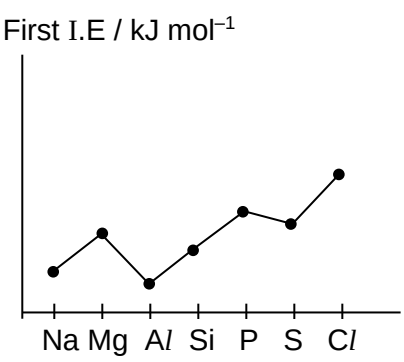
( )

5. Which of the following sketches show the correct trend in the stated property, for the elements in the third period of the Periodic Table?

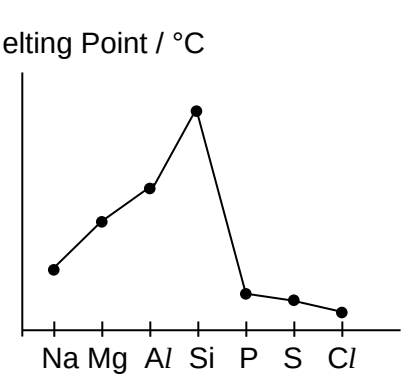
A



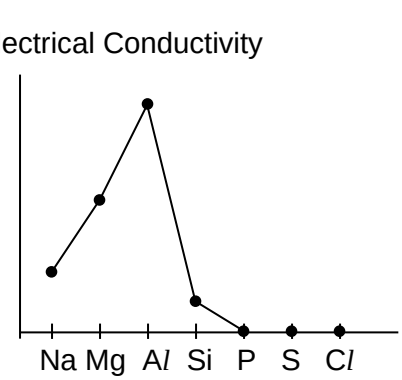
C



B



D



( )

Solutions to Checkpoint 1

Q1: D

- **Nuclear charge increases**.
- **Number of filled principal quantum shells and shielding effect remains the same** since the anions (P^{3-} to Cl^-) are **isoelectronic**.
- There are **stronger** (electrostatic) **forces of attraction** between the nucleus and the valence electrons.
- **Ionic radius decreases**.

Q2: A

- A **big decrease/drop in first ionisation energy of element** between **group 18 (period 2) and group 1 (period 3)**. The electron to be removed from group 1 is from an **outer principal quantum shell,** which is further away and more shielded from the nucleus, thus it experiences **weaker** (electrostatic) **forces of attraction** with the nucleus.
- **X is Mg** and the element on its right is **Al.**
  The first ionisation energy of Al is lower than that of Mg.
  $_{12}Mg : 1s^2\ 2s^2\ 2p^6\ 3s^2$ $_{13}Al : 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
  The **3*p*** electron to be removed from Al is at a **higher** energy level compared to the **3*s*** electron to be removed from Mg.

Less energy is required to overcome the **weaker** (electrostatic) **forces of attraction** between the nucleus and the valence **3*p*** electron in Al.

Q3: A

A: Electronegativity difference between the elements in each compound decreases.

	NaF	MgO	AlN	SiC
Electronegativity difference	3.98–0.93 =3.05	3.44–1.31 =2.13	3.04–1.61 =1.43	2.55–1.90 =0.65

Note: Calculations not required.

B: All the 4 compounds are iso–electronic with 20 electrons.

C: The bonding becomes increasingly covalent due to the **decreasing** electronegativity differences between the 2 elements in each compound.

D: Electrical conductivity **decreases** when dissolved in solution as the SiC is a covalent compound and will not produce mobile ions in water.

Q4: B

A: The atomic radius of Group 2 metals is **bigger** as it has smaller nucleus charge than Group 13 metals but experiences similar shielding effect.

B: Group 2 metals form bigger size cations and only contribute 2 valence electrons for delocalisation hence resulting in weaker (electrostatic) forces of attraction between cations and the sea of delocalised electrons. Therefore, lower melting points.

C: Group 2 metals have **higher** first ionisation energies than Group 13 metals.
(same explanation as Mg and Al. Refer to pg 11)

D: The interatomic distances in the metal lattice of Group 2 metals are **bigger** as they form bigger size atoms than Group 13.

Q5: D

A: The y–axis should be ionic radius.

B: The melting point of S_8 should be higher than P_4 as it has greater number of electrons hence, the strength of instantaneous dipole–induced dipole interactions is stronger.

C: The first I.E. of Al should not be lower than the first I.E. of Na.

D: Na, Mg and Al contains sea of delocalised electrons which acts as charge carrier. Si only conduct electricity at high temperature. P, S and Cl are covalent compounds hence non–conductor.

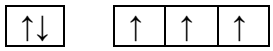
3. CHEMICAL PROPERTIES OF PERIOD 3 ELEMENTS AND THEIR COMPOUNDS

- The presence of 3d subshell in the valence shell of Period 3 elements enables Group 15 to 17 elements (P, S and C*l*) to **expand their octet structures** by excitation of valence electrons into **energetically accessible and available 3d orbitals.**

Example:

PCl_3

P:



3s 3p 3d

PCl_5

P:

3s 3p 3d

- The **maximum oxidation number** that the elements in Period 3 can attain has the **same numerical value** as their unit digit of the **group number**, which also corresponds to the number of valence electrons in the outermost principal quantum shell.

- Due to the *increasing electronegativity* across the Period, the structure of oxides and chlorides formed across the period changes from **giant ionic lattice structure to simple molecules**.
  The different structures of these compounds affect the type of bonding which in turn affect their physical and chemical properties. This will be covered in subsequent sections.

IMPORTANT:
Students must be able to write balanced chemical equations for all the reactions covered.

3.1 Oxides

Period 3 elements react with oxygen to form the following oxides:

*Oxides of chlorine are not in syllabus.

Elements	Na	Mg	A*l*	Si	P	S	C*l**
Formula of Oxide	Na_2O	MgO	Al_2O_3	SiO_2	(P_4O_6) P_4O_{10}	(SO_2) SO_3	(Cl_2O) Cl_2O_7
Oxidation number of element	+1	+2	+3	+4	(+3) +5	(+4) +6	(+1) +7

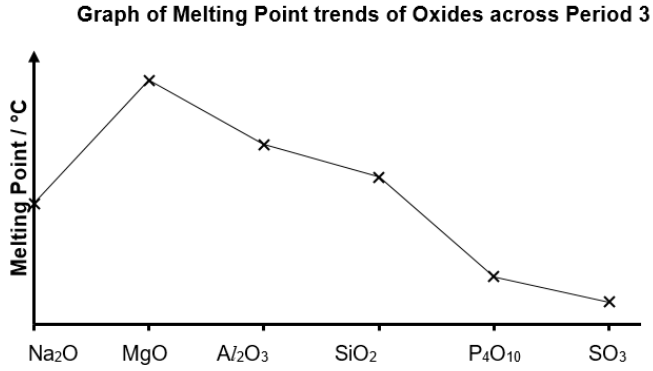
Why are the oxidation states of Period 3 elements in their oxides always positive?
Solution: Oxygen is more electronegative than all the period 3 elements.

3.1.1 Reaction with Oxygen (combustion) (SDL)

Scan the QR codes to view the observations. *You do not need to memorise colour of flame*.

Elements	Reaction with oxygen	Observations
Na	$2\ Na\ (s) + ½O_2\ (g) \rightarrow Na_2O\ (s)$ white solid	Burns readily in air or oxygen with a **brilliant yellow flame** to form an ionic oxide.
Mg	$Mg\ (s) + ½O_2\ (g) \rightarrow MgO\ (s)$ white solid	Burns with a **bright white light** to form MgO. The reaction produces a large amount of heat and light.
A*l*	$2\ Al\ (s) + {}^3/_2O_2\ (g) \rightarrow Al_2O_3\ (s)$ white solid	Easily oxidised in air to form Al_2O_3 that forms a <u>protective coating</u> and prevent the metal beneath from further attack by O_2 and water. When heated to about 800 ºC, A*l* burns with a **brilliant white flame**.
Si	$Si\ (s) + O_2\ (g) \rightarrow SiO_2\ (s)$	Burns with a **brilliant yellow white flame**.
P	<u>In limited oxygen</u>, $4P\ (s) + 3O_2\ (g) \rightarrow P_4O_6\ (s)$ <u>In excess oxygen</u>, $4P\ (s) + 5O_2\ (g) \rightarrow P_4O_{10}\ (s)$	Burns on heating with a **yellow flame**. *(P_4O_6 is oxidised to P_4O_{10} in air)*
S	$S\ (s) + O_2\ (g) \rightarrow SO_2\ (g)$ yellow solid <u>At 450 °C and V_2O_5 / Pt catalyst</u> $2SO_2\ (g) + O_2\ (g) \rightleftharpoons 2\ SO_3\ (g)$	Sulfur burns with a **blue flame** to form a colourless gas SO_2. In the presence of excess O_2 and V_2O_5 or Pt catalyst, 450 ºC, SO_2 can be further oxidised to SO_3.

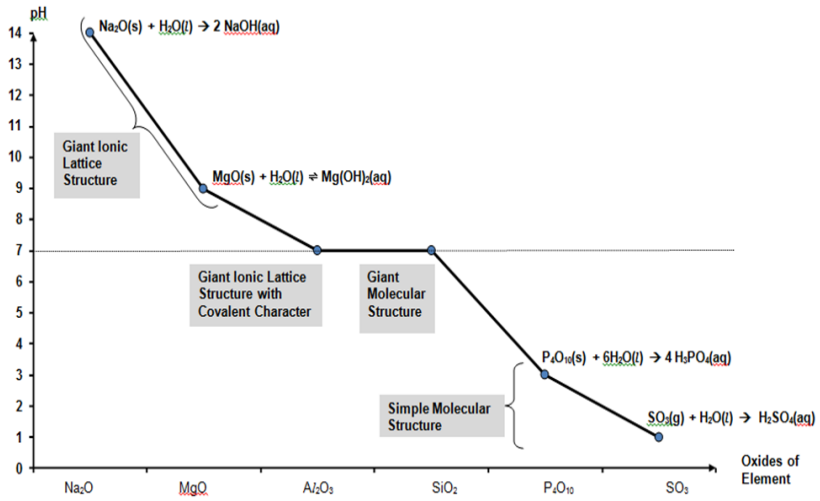
3.1.2 General Properties of Oxides



*P_4O_6 and SO_2 are not in syllabus.

Formula of Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10} (P_4O_6)	SO_3 (SO_2)
Structure/ Bonding	Giant ionic lattice structure		Giant ionic lattice structure with covalent character	Giant molecular structure	Simple molecular structure	
	Ionic bonds are formed due to **large difference in electronegativity**. This results in *electron transfer.*			Covalent bonds are formed due to **small difference in electronegativity**. This results in *electron sharing.*		
	Covalent character present in Al_2O_3 due to Al^{3+} having a **smaller cation size** and a **larger positive charge**. This results in Al^{3+} having a **high charge density**, causing the electron cloud of O^{2-} to become **more polarisable.**					
	Ionic character **decreases** → Covalent character **increases**					
Melting Point /°C	Very high			Very high	low	
	Sublimes at 1132	2852	2027	1610	340	17
	Refer to explanation of melting points in 'Chemical Bonding'.					
Electrical Conductivity	Does not conduct electricity in solid state. Good conductor in molten and aqueous state **Presence of mobile ions as mobile charge carriers** in the aqueous or molten state.			Non–conductor **Absence** of delocalised electrons or free mobile ions as **mobile charge carriers** in their molecular structures**.**		

3.1.3 Reactions of Oxides with Water



The **structure** of the oxides affects its nature.

Oxides with **giant ionic structure** are **basic** and react with acid.

Oxides with **molecular structure** are **acidic**.

Hence, Al_2O_3 which has **ionic structure with covalent characteristics** is **amphoteric** and reacts with both acid and alkalis.

Formula	Nature(1)	Solubility in water	Equation(s)	Resulting solution	Universal indicator
Na_2O	Basic	Soluble/ dissolves readily	$Na_2O(s) + H_2O(l) \rightarrow 2\ NaOH(aq)$	Strongly alkaline (pH ≈13)	violet
MgO		Sparingly soluble	$MgO(s) + H_2O(l) \rightleftharpoons Mg(OH)_2(aq)$	Weakly alkaline (pH ≈ 9)	blue
Al_2O_3	Amphoteric	Insoluble(1)	No reaction	pH 7(1)	green
SiO_2	Acidic	Insoluble(1)	No reaction	pH 7(1)	green
(P_4O_6) **P_4O_{10}**		Soluble/ dissolves readily	(2) $P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$ $P_4O_6(s) + 6H_2O(l) \rightarrow 4H_3PO_3(aq)$ (not required)	acidic (pH 2−3)	orange
(SO_2) **SO_3**		Soluble/ dissolves readily	$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$ $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$ (not required)	H_2SO_4 is a strong acid (pH 1) H_2SO_3 is a weak acid (pH 2–3)	red Orange

Important points to note:

(1) The insolubility in water is due to insufficient energy to compensate for energy required to overcome the very strong ionic bonds in Al_2O_3 or extensive network of strong covalent bonds in SiO_2. pH remains at 7 (neutral) as these oxides cannot dissolve.

(2) Reaction with water is not a redox reaction. Hence, the **oxidation state** of Period 3 element remains the **same** in both reactant and product.
Example: P_4O_{10} forms $H_3PO_4(aq)$ in water where P in both compounds has a +5 oxidation state.

Tip: When writing balanced equations, students can use this to deduce that the product of P_4O_{10} in water cannot be H_3PO_3. The P has a +5 oxidation state in P_4O_{10} and +3 oxidation state in H_3PO_3.

3.1.4 Reactions of Oxides / Hydroxides with Acids or Bases

Formula	Nature	Reaction	Equation(s)
Na_2O **NaOH**	basic	Undergoes **neutralisation with acids**	$Na_2O(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l)$ $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
MgO **$Mg(OH)_2$**			$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O\ (l)$ $Mg(OH)_2 + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O\ (l)$
Al_2O_3 **$Al(OH)_3$**	amphoteric	Undergoes **neutralisation with both acids and bases** to form salt and water	$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O\ (l)$ $Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2Na[Al(OH)_4]\ (aq)$ sodium aluminate $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O\ (l)$ $Al(OH)_3(s) + NaOH(aq) \rightarrow Na[Al(OH)_4]\ (aq)$
SiO_2	weakly acidic	Reaction requires heat and concentrated NaOH	Only with **strong heating** (400^oC) and **conc. base** $SiO_2(s) + 2NaOH(conc) \rightarrow Na_2SiO_3(aq) + H_2O(l)$ sodium silicate
(P_4O_6) **P_4O_{10}**	acidic	Readily undergoes **neutralisation with bases**	$P_4O_{10}(s) + 12NaOH(aq) \rightarrow 4Na_3PO_4(aq) + 6H_2O(l)$ $P_4O_6(s) + 8NaOH(aq) \rightarrow 4Na_2HPO_3(aq) + 2H_2O(l)$ (not required)
(SO_2) **SO_3**			$SO_3(g) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$ $SO_2(g) + 2NaOH(aq) \rightarrow Na_2SO_3\ (aq) + H_2O(l)$ (not required)

Note: P_4O_6 and SO_2 are not in the syllabus.

Writing chemical equations

Ability to write the chemical equations for all reactions discussed in this topic is critical. However, many students find it challenging to memorise the many chemical equations. Students can use the following concepts to help them memorise or balance these equations.

(1) Identify the reaction types e.g. neutralisation, combustion, oxidation.
(2) Identify the products formed in the reaction based on the type of reaction.
(3) For some aqueous systems, **AOHE** can be used to balance the equations.

Example:

<table>
<tr><td colspan="3">How to write the chemical equation for reaction of P_4O_{10} in NaOH?</td></tr>
<tr><td colspan="2">Identify the type of reaction</td><td>Neutralisation</td></tr>
<tr><td colspan="2">Identify the product.

Write the reactant and product in ionic form.</td><td>Since this is neutralisation, salt and water are formed. Since this is not a redox reaction, there is no change in the oxidation number of P, i.e. +5 in both reactant and product as well.

We can identify PO_4^{3-} as the product instead of PO_3^{3-}.

P_4O_{10} → PO_4^{3-}

Do not include H_2O at this stage as water will be balanced during O in AOHE.</td></tr>
<tr><td rowspan="5">Balance the atoms using AOHE.</td><td>A</td><td rowspan="5">P_4O_{10} → <u>4</u> PO_4^{3-}
<u>6 H_2O</u> + P_4O_{10} → 4 PO_4^{3-}
6 H_2O + P_4O_{10} → 4 PO_4^{3-} + <u>12 H^+</u>
Both sides are electrically neutral hence, electron is not required in this step. This is consistent that neutralisation is not a redox reaction and will not involve transfer of electrons.

<u>12 OH^-</u> + 6 H_2O + P_4O_{10} → 4 PO_4^{3-} + <u>12 OH^-</u> + <u>12 H^+</u>

<u>12 OH^-</u> + P_4O_{10} → 4 PO_4^{3-} + 6 H_2O

<u>12 NaOH</u> + P_4O_{10} → <u>4 Na_3PO_4</u> + 6 H_2O</td></tr>
<tr><td>O</td></tr>
<tr><td>H</td></tr>
<tr><td>E</td></tr>
<tr><td>Remove H^+ in alkaline condition</td></tr>
<tr><td>Add spectator ions</td></tr>
</table>

Exercise: Try writing equations for the following reactions.

		Equations
1	SO_3 and NaOH	
2	Al_2O_3 and NaOH	
3	SiO_2 and NaOH	

Checkpoint 2

4

1. Which compound is not a product of the reaction between an oxide of a third period element and water?

A NaOH
B H_2SiO_3
C H_3PO_4
D H_2SO_4

( )

2. The properties of the oxides of four Period 3 elements **N, P, Q** and **R** are given below.

The oxide of **N** is insoluble in water and in dilute acid but soluble in concentrated sodium hydroxide. The oxide of **P** is amphoteric. The oxide of **Q** reacts with dilute sodium hydroxide at room temperature. The oxide of **R** dissolves in water to form a strong alkaline solution.

Which of the following is correct in order of increasing proton number?

A **R, P, N, Q**
B **N, P, Q, R**
C **R, N, P, Q**
D **R, Q, N, P**

( )

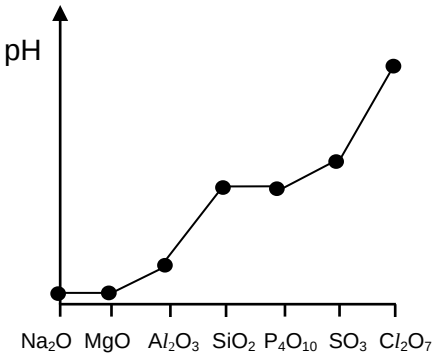
3. **P**, **Q** and **R** are main group elements in the same period of the Periodic Table.
The oxide of **P** is basic, the oxide of **Q** is acidic and the oxide of **R** is amphoteric.
What is the order of increasing atomic number for these elements, starting from the lowest?

A **P** < **Q** < **R**
B **P** < **R** < **Q**
C **Q** < **R** < **P**
D **R** < **P** < **Q**

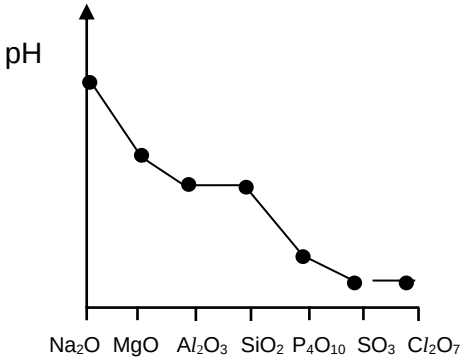
( )

4. The highest oxides of the element sodium to chlorine are separately added to water. Which diagram best represents the pH of the solutions produced?

A



B



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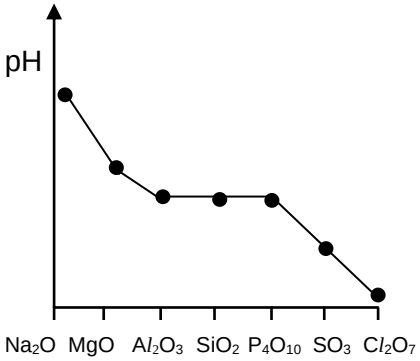

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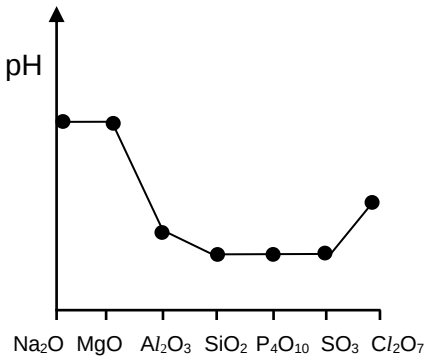

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Solutions to Checkpoint 2

Q1: B
SiO_2 is insoluble / does not react with water due to the extensive network of strong covalent bonds between the Si and O atoms in giant molecular structure.

Q2: A
N: SiO_2 is insoluble due to extensive network of strong covalent bonds between atoms
P: Al_2O_3 is amphoteric.
Q: P_4O_{10} is acidic.
R: Na_2O is basic.

Option A (Na_2O, Al_2O_3, SiO_2, P_4O_{10}) is in order of increasing proton number.

Q3: B
P: Na_2O, MgO
R: Al_2O_3
Q: SiO_2, P_4O_{10} , SO_3

Option B (P < R < Q) is in order of increasing atomic number.

Q4: B

	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
pH (oxide with water)	13	9	7	7	2–3	2–3	2

Al_2O_3 and SiO_2 are both insoluble in water hence pH is 7.

3.2 Chlorides

The following **chlorides** of Period 3 elements are formed:

* PCl_3 is not in syllabus.

Elements	Na	Mg	Al	Si	P
Formula of Chloride	NaCl	$MgCl_2$	$AlCl_3$ Al_2Cl_6	$SiCl_4$	(PCl_3) PCl_5
Oxidation number of element	+1	+2	+3	+4	(+3) +5

Note: Chlorides of sulfur is not in the syllabus.

- The **maximum oxidation number** that the elements in Period 3 can attain has the **same numerical value** as their unit digit of the **group number**, which also corresponds to the number of valence electrons in the outermost principal quantum shell.

Oxidation States

Elements in Groups 15 to 16 can also exhibit multiple oxidation states because they can *expand their octet structures* by excitation of valence electrons into *energetically accessible and available 3d orbitals.*

Why are the oxidation states of Period 3 elements in their chlorides always positive?
Solution: Chorine is more electronegative than all the period 3 elements.

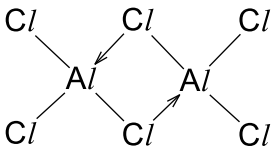
3.2.1 **Reaction with Chlorine** (SDL)

Scan QR codes to view the reactions. Observe the **colour**, **states** and **relative rate of the reactions** in these videos.

Elements	Reaction with chlorine	Observations
Na	$2Na\ (s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$	Very vigorous
Mg	$Mg\ (s) + Cl_2\ (g) \rightarrow MgCl_2\ (s)$	Vigorous
Al	$2Al\ (s) + 3Cl_2\ (g) \rightarrow 2AlCl_3\ (s)$	Vigorous
Si	$Si\ (s) + 2Cl_2\ (g) \rightarrow SiCl_4\ (l)$	Slow
P	$2\ P\ (s) + 3Cl_2\ (g) \rightarrow 2PCl_3\ (l)$ $PCl_3\ (l) + Cl_2\ (g) \rightarrow PCl_5\ (s)$ With excess Cl_2 (g), PCl_3 is oxidised to PCl_5.	Slow

3.2.2 General Properties of Chlorides

Some general properties of the chlorides are shown below:

<table>
<tr><th>Formula of Chloride</th><th>NaCl</th><th>MgCl$_2$</th><th>AlCl$_3$
Al$_2$Cl$_6$</th><th>SiCl$_4$</th><th>PCl$_3$
PCl$_5$</th></tr>
<tr><td rowspan="3">Structure/ Bonding</td><td colspan="2">Giant ionic lattice Structure</td><td colspan="3">Simple molecular structure</td></tr>
<tr><td colspan="5">The trend arises due to the decreasing difference in electronegativity between period 3 element and chlorine across the period.</td></tr>
<tr><td colspan="5">Ionic character decreases
→
Covalent character increases</td></tr>
<tr><td>Melting Point /°C</td><td>808</td><td>714</td><td>192</td><td>–70</td><td>PCl$_3$ (–92)
PCl$_5$ (167)</td></tr>
<tr><td>Boiling Point /°C</td><td>1465</td><td>1418</td><td>AlCl$_3$ sublimes and dimerises to form Al$_2$Cl$_6$ at 180 ^{0}C
 
Al$_2$Cl$_6$ breaks up to form AlCl$_3$ at 800 ^{0}C</td><td>57</td><td>PCl$_3$ (76.1)

PCl$_5$ sublimes at 160 ^{0}C</td></tr>
<tr><td></td><td colspan="5">Refer to explanation of melting and boiling points in ‘Chemical Bonding’.</td></tr>
<tr><td>Electrical Conductivity</td><td colspan="2">Non–conductor in solid state but are good conductor in molten and in aqueous state

Presence of mobile ions as mobile charge carriers to conduct electricity.</td><td colspan="3">Non–conductor in all states (with the <u>exception</u> of AlCl$_3$ in aqueous state)

Absence of mobile charge carriers because electrons are localised in covalent bonds and not mobile to conduct electricity.</td></tr>
</table>

3.2.3 Reactions of Chlorides with Water

The nature of the chlorides changes from **neutral to acidic** across the period**.**

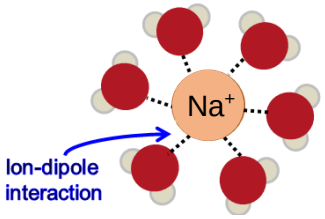
When chlorides interact with water, *hydration* and *hydrolysis* may take place.

Metallic chlorides in water

Hydration

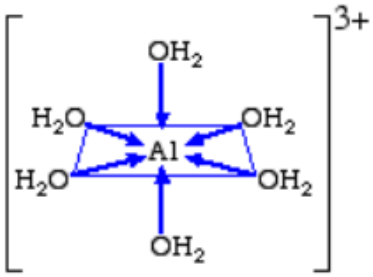
- Hydration is a physical process.
- Water molecules surround the **cations**, forming **ion–dipole interactions** with **e.g. $Na^+(aq)$**

E.g. $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$



and **dative bonds** in **$[Mg(H_2O)_6]^{2+}$ and $[Al(H_2O)_6]^{3+}$**

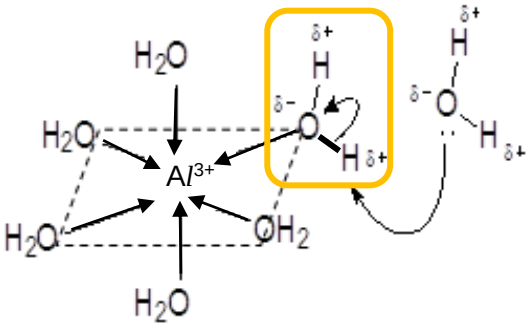
E.g. $AlCl_3(s) + 6H_2O(l) \rightarrow [Al(H_2O)_6]^{3+}(aq) + 3Cl^-(aq)$



Hydrolysis

- Hydrolysis involves the **breaking of a covalent bond** in a molecule using water.
- The reaction can occur between an ion/a molecule and water molecules.
- Hydrolysis **changes the pH** of a solution.
- Cations with **high charge density like Al^{3+}, Cr^{3+} and Fe^{3+}** undergo hydrolysis.
- Na^+ has low charge density and does not hydrolyse in water.

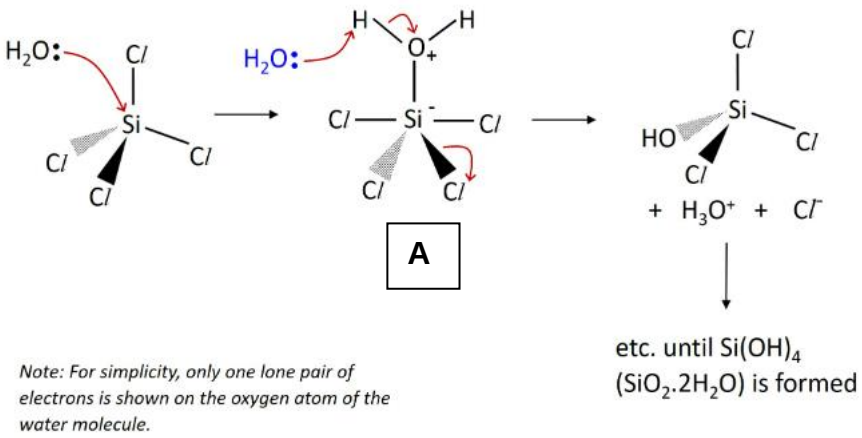
E.g. $[Al(H_2O)_6]^{3+}(aq) + H_2O(l) \rightleftharpoons [Al(OH)(H_2O)_5]^{2+}(aq) + H_3O^+(aq)$



1) High charge density of Al^{3+}
2) Greater polarisation of water molecule
3) O–H bond weakens and breaks
4) H^+ is released into water

Non–metallic chlorides in water

E.g. $SiCl_4$



Note: For simplicity, only one lone pair of electrons is shown on the oxygen atom of the water molecule.

$SiCl_4$ is able to undergo **hydrolysis** because Si has **available** and **energetically accessible 3d orbitals** that allow H_2O to form a **dative bond** with Si, bringing H_2O to close proximity for reaction.

Substance **A** has 5 bond pairs surrounding Si hence expanded octet which is possible for Period 3 elements.

Note: Writing this mechanism is not required.

Why is CCl_4 unable to hydrolyse in water?

- Although carbon is in the same group as silicon, CCl_4 does not undergo hydrolysis in water because of the much **smaller size of the C atom relative to the large Cl atoms** which **hinders** the approach of the water molecules.

  In addition, C does **not** have **available and energetically accessible 3d orbitals** to accommodate the lone pair of electrons from water unlike Si.

Reactions of Chlorides with Water

Formula of Chloride	NaCl	$MgCl_2$	$AlCl_3$ Al_2Cl_6	$SiCl_4$	PCl_3 PCl_5
Reaction with H_2O	Hydration only	Hydration and **slight** hydrolysis	Hydration and **appreciable** hydrolysis	Hydrolysis only	
pH	Neutral (pH 7)	Slightly acidic (≈ pH 6.5)	Acidic (pH 3)	Strongly acidic (pH 2) due to formation of HCl	
Reason	Due to **low charge density** of Na^+ and Cl^-, Na^+ and Cl **will not hydrolyse** in water thus forming a neutral solution.	Due to the **smaller cation size** and a **larger positive charge,** and hence a **higher charge density** of Mg^{2+} (higher polarising power), $MgCl_2$ is an **ionic chloride with some covalent character**. After dissolving, the hydrated magnesium ion undergoes **slight hydrolysis** to form a slightly acidic solution of pH 6.5.	In a large quantity of water, hydrated ions are formed. Due to the **smaller cation size** and a **larger positive charge,** and hence a **higher charge density** of Al^{3+} (high polarising power), it is able to draw electrons to itself from the oxygen atoms of the neighbouring water molecules, which further polarises the O–H bonds, thereby producing H^+ in the solution. Hence, $AlCl_3$ **hydrolyses** in water to form an acidic solution of pH 3. When limited amount of water is used, a white $Al(OH)_3$ solid and white HCl fumes are formed instead.	$SiCl_4$, PCl_5 and all other covalent chlorides undergo **complete hydrolysis in water** to produce very acidic solutions of pH 1–2. View the video for reaction between $SiCl_4$ and H_2O.	
Equation (Excess H_2O)	$NaCl\,(s) + aq \rightarrow Na^+\,(aq) + Cl^-\,(aq)$	$MgCl_2\,(s) + 6H_2O\,(l) \rightarrow [Mg(H_2O)_6]^{2+}\,(aq) + 2Cl^-\,(aq)$ $[Mg(H_2O)_6]^{2+}\,(aq) \rightleftharpoons [Mg(H_2O)_5(OH)]^+\,(aq) + H^+\,(aq)$	$AlCl_3\,(s) + 6\,H_2O\,(l) \rightarrow [Al(H_2O)_6]^{3+}(aq) + 3\,Cl^-\,(aq)$ $[Al(H_2O)_6]^{3+}\,(aq) \rightleftharpoons [Al(H_2O)_5(OH)]^{2+}\,(aq) + H^+\,(aq)$	$SiCl_4\,(l) + 4\,H_2O\,(l) \rightarrow SiO_2.2H_2O\,(s) + 4\,HCl\,(aq)$ $PCl_3\,(l) + 3\,H_2O\,(l) \rightarrow H_3PO_3\,(aq) + 3\,HCl\,(aq)$ * PCl_3 is not in syllabus. $PCl_5\,(s) + 4\,H_2O\,(l) \rightarrow H_3PO_4\,(aq) + 5\,HCl\,(aq)$	
Equation (Limited H_2O)	–	–	$AlCl_3\,(s) + 3H_2O\,(l) \rightarrow Al(OH)_3\,(s) + 3HCl\,(g)$ white solid white fumes	$PCl_5\,(s) + H_2O\,(l) \rightarrow POCl_3\,(l) + 2\,HCl(g)$	

Checkpoint 3

4

1. Which one of the following is a correct statement about aluminium?

A It forms amphoteric chlorides.
B It forms an oxide that is insoluble in water.
C It forms a solid oxide that sublimes at 180 ^{o}C.
D It forms oxides that can exist as dimers in the vapour phase.

( )

2. The labels had become detached from four bottles in the laboratory. A student realised that the contents of one of them could easily be identified because, on addition of water it would **not** give fumes that turns moist blue litmus paper red.

Which compound would **not** give out fumes that turn moist blue litmus paper red?

A Al_2Cl_6
B $MgCl_2$
C PCl_5
D $SiCl_4$

( )

3. Aluminium chloride catalyses certain reactions by forming carbocations (carbonium ions) with chloroalkanes as shown in the following equation.

$$RCl + AlCl_3 \rightarrow R^+ + AlCl_4^-$$

This can occur because

A $AlCl_3$ is a covalent molecule.
B $AlCl_3$ exist as the dimer Al_2Cl_6 in the vapour.
C The chlorine atom in R*Cl* has a vacant p orbital.
D The aluminium atom in $AlCl_3$ has an incomplete octet of electrons.

( )

4. Element X is a solid with a very low electrical conductivity at room temperature. It forms only one chloride, which is a liquid at room temperature and is a non–conductor of electricity. The chloride hydrolyses in water forming a white solid and a strongly acidic solution.

Which of the following could be X?

A Aluminium
B Phosphorus
C Silicon
D Sulfur

( )

Solutions to Checkpoint 3

Q1 B

A: AlCl_3 is acidic in nature.

B: Al_2O_3 has high lattice energy hence insoluble in water.

C: AlCl_3 (not Al_2O_3) sublimes at 180 ^{0}C.

D: AlCl_3 (not Al_2O_3) dimerises to form Al_2Cl_6 at 180 ^{0}C

Q2 B

MgCl_2 (s) + 6 H_2O (l) → $[Mg(H_2O)_6]^{2+}$ (aq) + 2 Cl^- (aq)

$[Mg(H_2O)_6]^{2+}$ (aq) ⇌ $[Mg(H_2O)_5(OH)]^+$ (aq) + H^+ (aq)

No fumes given out.

Al_2Cl_6, PCl_5 and SiCl_4 give out white HCl fumes which are acidic.

Q3 D

In AlCl_3 molecule, Al atom only has 6 outermost electrons hence it is able to accept the lone pair of electrons from Cl^- ion to form complete octet structure.

Q4 C

Si is a semiconductor which can only give rise to delocalised electrons at high temperatures.

SiCl_4 (l) has a simple molecular structure hence is a non–conductor of electricity.

SiCl_4 (l) + 4 H_2O (l) → $SiO_2.2H_2O$ (s) + 4 HCl (aq)

(white solid) (strongly acidic solution)

4. THE GROUP 2 ELEMENTS

4.1 Introduction

- Group 2 elements (alkaline earth metals) are one of the two groups of *s*−block elements. Their outer electrons are in *s* orbitals, with a valence shell configuration, **ns^2**.

- Group 2 elements are reactive metals. They occur naturally in combination with other elements in mineral deposits and seawater as salts, and are extracted through electrolysis.

- In their pure form, the elements exist as **silvery grey metals** which tarnish readily in air.

- Group 2 elements are all **good reducing agents** and only show **one oxidation state of +2** in their compounds

- Group 2 elements show only the +2 oxidation state because:
  - Group 2 metals have generally low first and second ionisation energies.
  - A small amount of energy is required to remove the two outer *s* electrons as they experience weak electrostatic forces of attraction due to shielding effect.
  - Subsequent electrons involve the loss of inner shell electrons, which require more energy, hence higher oxidation states are not favourable.

- Most Group 2 compounds are **ionic** (Be is an exception), containing **M^{2+}** ions.

- All Group 2 oxides formed are **white** and have **extremely high melting points**.

4.2 Physical properties of Group 2 elements

- Several physical properties of Group 2 elements and their trends are summarised in Table 1.1

Table 1.1 Physical Properties of Group 2 Metals

	Z	electronic structure	atomic radius / nm	ionic radius / nm	1^{st} + 2^{nd} I.E. / kJ mol^{-1}	electro−negativity	m.p. / °C	b.p. / °C
Be Beryllium	4	[He] $2s^2$	0.112	0.031	2660	1.5	1280	2770
Mg Magnesium	12	[Ne] $3s^2$	0.160	0.065	2186	1.2	650	1110
Ca Calcium	20	[Ar] $4s^2$	0.197	0.099	1740	1.0	840	1440
Sr Strontium	38	[Kr] $5s^2$	0.215	0.113	1608	1.0	768	1380
Ba Barium	56	[Xe] $6s^2$	0.217	0.135	1468	0.9	714	1640

4.3 Chemical properties of Group 2 elements

4.3.1 Reducing Power

Group 2 elements are strong reducing agents.

- The reducing power of the metals can be measured by their ***standard electrode potential***.

Down the group,

- the standard electrode potential $E^{\ominus}$ values become less positive,
- the reducing power of the metals increases.
- the metals have higher tendency to be oxidised. (i.e.: The metal atoms lose their valence electrons **more readily**.

Down Group 2, metal atoms lose their valence electrons **more readily** because

- The number of filled principal quantum shells **increases**.
- Valence electrons are **further** and **more shielded** from the nucleus.
- There are **weaker** electrostatic forces of attraction between the nucleus and the valence electrons.

	$E^{\ominus}$ / V
$Be^{2+} + 2e \rightleftharpoons Be$	–1.85
$Mg^{2+} + 2e \rightleftharpoons Mg$	–2.38
$Ca^{2+} + 2e \rightleftharpoons Ca$	–2.87
$Sr^{2+} + 2e \rightleftharpoons Sr$	–2.89
$Ba^{2+} + 2e \rightleftharpoons Ba$	–2.90

Standard Electrode Potential, $E^{\ominus}$

$M^{2+} + 2e^{-} \rightleftharpoons M \qquad E^{\ominus}$ /V

- In the Data Booklet, half−equations are represented with reversible arrows.
  - The **forward** reaction is the **reduction** of M^{2+}.
  - The **backward** reaction is the **oxidation** of M.

- A **positive** $E^{\ominus}$ value indicates that the **reduction** reaction is predominant. The **more positive** the $E^{\ominus}$ value,
  - the **greater the tendency** for **reduction of M^{2+}** to occur
  - the **stronger** the **oxidising power of M^{2+}**

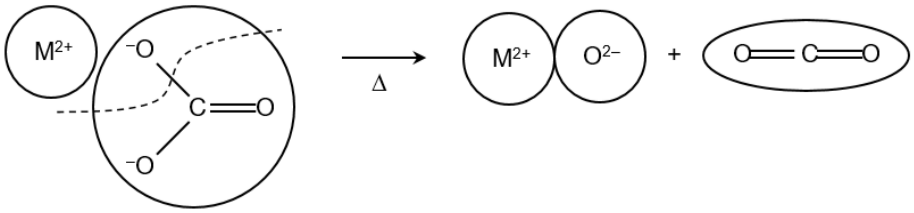
- A **less positive** $E^{\ominus}$ value indicates that the **oxidation** reaction is predominant. The **less positive** the $E^{\ominus}$ value,
  - the **greater the tendency** for **oxidation of M** to occur
  - the **stronger** the **reducing power of M**

Note: The concept of 'Standard Electrode Potential' was covered in greater details in the topic 'Electrochemistry'.

4.3.2 Thermal Decomposition of Group 2 Carbonates

Group 2 carbonates can be broken down by heating into their oxides and carbon dioxide.

Equation: $MCO_3(s) \rightarrow MO(s) + CO_2(g)$



The following table shows the decomposition temperature of the Group 2 carbonates.

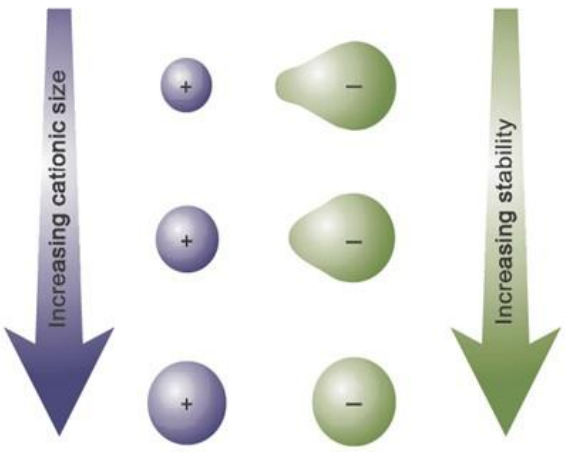
M	Temperature of decomposition / °C	Thermal Stability
Be	250	INCREASES
Mg	540	
Ca	900	
Sr	1289	
Ba	1360	

From the above table, down **Group 2**:

- The **ease** of thermal decomposition of carbonates **decreases**.
- **Thermal stability** of carbonates **increases**.

Explanation:
Down the group,

- **ionic radius** of M^{2+} **increases** so the **charge density** of the M^{2+} **decreases**.
- M^{2+} becomes less polarising.
- The electron cloud of CO_3^{2-} anion is less distorted so the **C–O covalent bond** within the CO_3^{2-} anion is **less weakened** down the group.
- **Ease of decomposition decreases** down the group OR
- **Thermal stability of carbonates increases** down the group.



What about the decomposition of Group 2 ***nitrates****?*

$$\mathbf{2M(NO_3)_2\,(s) \rightarrow 2MO\,(s) + 4NO_2\,(g) + O_2(g)}$$

The argument is exactly the same here. The small positive ions at the top of the Group polarise the nitrate ions (NO_3^-) more than the larger positive ions at the bottom.

Recall

General Terms:

- **Polarisation** is the distortion of electron cloud of an anion by a cation
- **Polarising power** of a cation is its ability to distort the electron cloud of an anion.
- **Polarisability** of an anion is the ease with which its electron cloud gets distorted

For a cation, polarising power $\propto$ charge density ($= \frac{\text{ionic charge}}{\text{ionic radius}}$)

- A small and highly charged cation has a high charge density, hence high polarising power, e.g., Mg^{2+} ion is more polarising than Ba^{2+} ion.

For an anion, polarisability depends on its charge and size.

- A large and highly charged anion has a high polarisability as the electron cloud is more diffused (or loosely held).
  e.g., CO_3^{2-} ion is more polarisable than O^{2-} ion.

Summary of Chemical Properties of Group 2 Elements and their Oxides & Carbonates

		Mg	Ca	Sr	Ba
Electronic Configuration		$[Ne]\ 3s^2$	$[Ar]\ 4s^2$	$[Kr]\ 5s^2$	$[Xe]\ 6s^2$
$E^{\ominus}(M^{2+}/M)$ / V		−2.38	−2.87	−2.89	−2.90
Reducing power		increases down the group →			
Reaction with O_2 [Redox]		$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$	$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$	$2Sr(s) + O_2(g) \rightarrow 2SrO(s)$	$2Ba(s) + O_2(g) \rightarrow 2BaO(s)$
Reaction with H_2O [Redox]	**Reactivity**	• reactivity increases down the group due to the **increase in reducing power** of the metals. →			
		• reacts quickly with steam • reacts very slowly in cold water to give a weakly alkaline solution	• reacts with cold water with increasing vigour to form alkaline solutions of increasing strength		
	Equation	<u>With steam</u> $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$ <u>With cold water</u> (very slow) $Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)$	<u>With cold water (slow)</u> $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$	<u>With cold water</u> $Sr(s) + 2H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$	<u>With cold water (very vigorous)</u> $Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$
Solubility of hydroxides		sparingly soluble (white suspension)	slightly soluble	soluble	soluble
pH of resulting solution		$Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$			
		~9	10 – 13		
		MgO	**CaO**	**SrO**	**BaO**
Reaction of oxides with water [Hydrolysis]		$MgO(s) + H_2O(l) \rightleftharpoons Mg(OH)_2(s)$	$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$	$SrO(s) + H_2O(l) \rightarrow Sr(OH)_2(aq)$	$BaO(s) + H_2O(l) \rightarrow Ba(OH)_2(aq)$
Trend of pH of resulting solution		$Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$			
Reaction of oxides with acid		$MO(s) + 2H^+(aq) \rightarrow M^{2+}(aq) + H_2O\ (l)$			
Solubility of oxides		sparingly soluble	slightly soluble	soluble	soluble
		$MgCO_3$	**$CaCO_3$**	**$SrCO_3$**	**$BaCO_3$**
Thermal stability	**Trend**	thermal stability of Group 2 carbonates <u>increases</u> down the group →			
		Reason: Down the group, − <u>ionic radius</u> of M^{2+} <u>increases</u>, hence, the <u>charge density</u> of the M^{2+} <u>decreases</u>, the M^{2+} becomes less polarising. − The electron cloud of CO_3^{2-} anion is less distorted. The <u>C−O covalent bond</u> within the CO_3^{2-} anions <u>is less weakened</u> down the group. − <u>Ease of decomposition decreases</u> down the group (i.e. <u>thermal stability of carbonates increases</u> down the group.)			
	Equation	$MCO_3(s) \rightarrow MO(s) + CO_2(g)$			

5 REAL WORLD APPLICATIONS OF GROUP 2 COMPOUNDS (SDL)

- **Magnesium oxide** (MgO) is used as a **refractory lining** in **high temperature furnaces** because of its high melting point and low reactivity.

- **Calcium carbonate** ($CaCO_3$), also known as **limestone**, is used in **construction materials**.

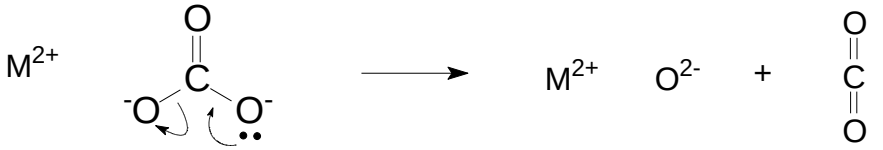
- **Calcium oxide** (CaO), also known as **quicklime**, is used in **agriculture** to overcome the problem of acidity in soil:
  - The pH of soil affects the ability of crops to absorb nutrients.
  - The optimum pH for most crops lies in the range of 6 to 7.
  - **Liming** (i.e. the process of adding basic calcium oxide to acidic soil) is employed to **reduce acidity** of soil in order to maximise crop yields.

- **Barium sulfate** ($BaSO_4$) is used in **'barium meal' test** in x−ray diagnostic work.

Additional Reading

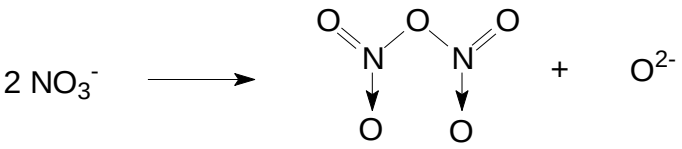
(a) Mechanism for thermal decomposition of MCO_3 (Reference: N2012/II/Q3)

Overall equation: $MCO_3(s) \rightarrow MO(s) + CO_2(g)$

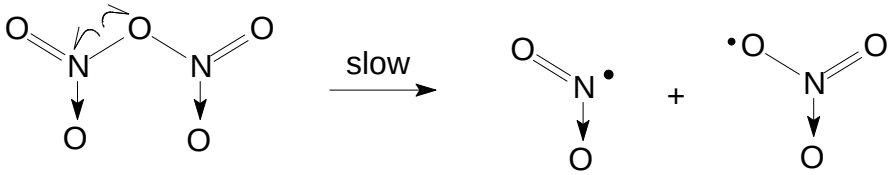


(b) Mechanism for thermal decomposition of $M(NO_3)_2$ (Reference: N2012/II/Q3)

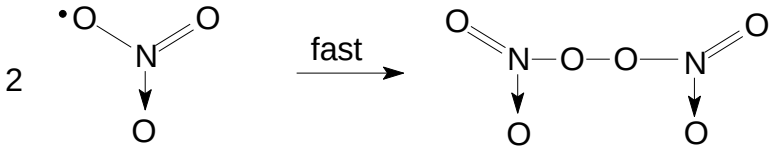
Overall equation: $M(NO_3)_2(s) \rightarrow MO(s) + 2\ NO_2(g) + \frac{1}{2}\ O_2(g)$



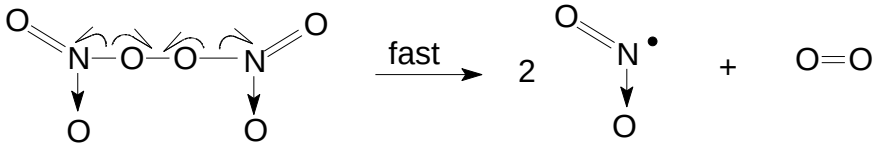
There is an initial homolytic breaking of an N–O bond to give $^{\bullet}NO_2$ and $NO_3^{\bullet}$ radicals



Dimerisation of the $NO_3^{\bullet}$ radicals to give N_2O_6



Dissociation of the dimer via further N–O bond cleavage to form the products.



6. THE GROUP 17 ELEMENTS

6.1 Introduction

- Group 17 elements (also called "halogens") are *p–block* elements.
- When halogens react with the other elements, they can complete their octet by **gaining an additional electron to form singly charged halide ions**, X^-, (oxidation state = –1).
- Halogens can also **form covalent bonds with another halogen atom,** example I–C*l*. Oxidation state of the more electronegative C*l* atom is assigned –1 and the less electronegative I atom = +1 in I–C*l*.
- All halogens are toxic. They have a pungent odour and irritate the skin.
  - Bromine causes burns that are very slow to heal.
  - Astatine (At) is radioactive and its most stable isotope has a half–life of 8.3 hours.

6.2 Physical Properties

Element	**Fluorine (F_2)**	**Chlorine (Cl_2)**	**Bromine (Br_2)**	**Iodine (I_2)**
Electronic Configuration Outer–shell configuration **$ns^2\ np^5$**	[He] $2s^2\ 2p^5$	[Ne] $3s^2\ 3p^5$	[Ar] $3d^{10}\ 4s^2\ 4p^5$	[Kr] $4d^{10}\ 5s^2\ 5p^5$
*Colour and physical state (r.t.p)	pale yellow gas	greenish yellow gas	reddish–brown liquid	black solid
*Colour in non–polar solvent (e.g. hexane)	very pale yellow	pale yellow	orange–red	purple
*Colour in aqueous solution	–	pale yellow	orange	brown (due to I_3^-)
Melting Point / °C	*Increases* down the group due to *increasing electron cloud size* that leads to *stronger instantaneous dipole–induced dipole attraction*.			
Boiling Point / °C				
#Bond energy / kJ mol^{-1}	158	244	193	151

Note:

- *Intensity of the colour of the halogens **increases** down the group.
- #F–F bond is weaker than C*l*–C*l* bond due to **repulsion between lone pairs of electron** on the 2 fluorine atoms which are **very near** to each other. This result in **less effective overlap** in the **orbitals** formed between F atoms compared to C*l* atoms.

6.3 Variation in Volatility

- **Volatility** is a measure of how readily a substance vaporises or changes from a liquid phase to a gas phase.
- Volatility of the halogens **decreases** down the group.
- Boiling point of the halogens **increases** down the group.

$$\text{volatility} \propto \frac{1}{\text{boiling point}}$$

Explanation:
- Halogens have **simple molecular structures** and are **non–polar**.
- Down the group, the **number of electrons** in the halogen molecules **increases**.
- **More energy** is required to overcome the **stronger instantaneous dipole–induced dipole attractions** between the halogen molecules that is further down the group.

6.4 Oxidising Strength of Halogens

- Halogens are **strong oxidising agents**.

  Given: $X_2 + 2e \longrightarrow 2X^-$

  When Group 17 halogen, X_2 gain electrons to form X^-, X_2 is reduced; it acts as an oxidising agent.
- Down the group, **oxidising power** of the halogens **decreases**.

Order of oxidising power: F_2 > Cl_2 > Br_2 > I_2

Explanation:
Down Group 17,
- The number of filled **principal quantum shells increases.**
- Valence electrons are **further and more shielded from the nucleus**.
- There are **weaker electrostatic forces of attraction between the nucleus and the valence electrons**.
- Thus, ability of X to gain electrons (get reduced) decreases, hence **oxidising power decreases.**

Examples of Halogens as Oxidising Agents

- Displacement reaction
  Stronger oxidising agent higher in the Group oxidises (and hence displaces) the halide ions in aqueous solution further down the Group.

$$F_2(aq) + 2Cl^-(aq) \rightarrow 2F^-(aq) + Cl_2(aq) \quad E_{cell} = +2.87 - (+1.36) > 0V$$
$$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq) \quad E_{cell} = +1.36 - (+0.54) > 0V$$

7 THE GROUP 17 HYDRIDES

7.1 Melting and Boiling Points of Group 17 hydrides (hydrogen halides)

hydrogen halide	melting point / ^{o}C	boiling point / ^{o}C
HF	–83.1	+19.5
HC*l*	–115	–85.0
HBr	–89.0	–67.0
HI	–51.0	–35.0

- Down the group, melting / boiling point of the hydrogen halides generally **increases**.

Explanation:
- Hydrogen halides have **simple molecular structures**.
- Down the group, the **number of electrons** in the hydrogen halide molecules **increases**.
- **Strength** of **instantaneous dipole–induced dipole attractions** between the hydrogen halide molecules **increases**
- **Larger amount of energy** is required to overcome the **stronger** instantaneous dipole–induced dipole attractions between the hydrogen halide molecules.

Note:
- HF has abnormally high melting and boiling points due to more energy required to overcome the **stronger hydrogen bonding** between HF molecules.

7.2 Thermal Stability of Group 17 hydrides (hydrogen halides)

hydrogen halide	**bond energy / kJ mol^{-1}**	**observation upon heating**	**equation**
HF	+562	do not decompose even upon strong heating	–
HC*l*	+431		
HBr	+366	strong heating yields reddish–brown fumes of Br_2	$2HBr(g) \rightarrow H_2(g) + Br_2(g)$
HI	+299	Purple fumes of I_2 observed when hot rod is plunged into a jar of HI, deposit of black solid I_2 obtained when cooled.	$2HI(g) \rightarrow H_2(g) + I_2(g)$

- Down the group, **thermal stability** of the hydrogen halides **decreases**.

Explanation:
- Down the group from F to I,
- **Valence orbital** of halogen used in bonding with hydrogen is **bigger** and **more** diffused
- **Overlap of orbitals between the H atom and halogen atom** is **less** effective
- **Bond energy**: H–F > H–C*l* > H–Br > H–I
- **Thermal stability**: HF > HC*l* > HBr > HI

Summary of Chemical Properties of Group 17 Elements (X_2) and Hydrogen Halides (HX)

<table>
<tr><th colspan="2"></th><th>F_2</th><th>Cl_2</th><th>Br_2</th><th>I_2</th></tr>
<tr><td colspan="2">Electronic Configuration</td><td>[He] $2s^2 2p^5$</td><td>[Ne] $3s^2\ 3p^5$</td><td>[Ar] $4s^2\ 4p^5$</td><td>[Kr] $5s^2\ 5p^5$</td></tr>
<tr><td colspan="2">$E^{\ominus}(X_2 / X^-)$ / V</td><td>+2.87</td><td>+1.36</td><td>+1.07</td><td>+0.54</td></tr>
<tr><td colspan="2">Oxidising power</td><td colspan="4">decreases down the group →
• Down the group, $E^{\theta}_{X_2 / X^-}$ becomes <u>less positive</u>.
⇒ lower tendency for X_2 to be reduced to X^-.
⇒ <u>oxidising power of halogens decreases</u> from F_2 to I_2.</td></tr>
<tr><td colspan="2">E.g. Displacement reaction</td><td colspan="4">• A more reactive (i.e. more strongly oxidising) halogen <u>displaces</u> a <u>less reactive</u> one <u>from its compound</u>:
• Equation: $X_2 + 2Y^- \rightarrow Y_2 + 2X^-$ (i.e: $F_2(aq) + 2Cl^-(aq) \rightarrow 2F^-(aq) + Cl_2(aq)$)</td></tr>
<tr><td colspan="2">E.g. Reaction with $S_2O_3^{2-}$</td><td colspan="4">• The strength of halogens as oxidising agents can be demonstrated as:
• Cl_2 and Br_2 oxidises $S_2O_3^{2-}$ to SO_4^{2-} [Change in Oxidation No of S : +2 to +6]
• while I_2 oxidises $S_2O_3^{2-}$ to $S_4O_6^{2-}$ [Change in Oxidation No of S : +2 to +2.5]</td></tr>
<tr><td rowspan="2">E.g. Reaction with H_2</td><td>Reactivity</td><td colspan="4">Reaction become less vigorous down the group →</td></tr>
<tr><td>Equation</td><td>$F_2 + H_2 \rightarrow 2HF$
Very rapid reaction</td><td>$Cl_2 + H_2 \rightarrow 2HCl$
Rapid reaction</td><td>$Br_2 + H_2 \rightarrow 2HBr$
Slow reaction</td><td>$I_2 + H_2 \rightleftharpoons 2HI$
No reaction unless strongly heated.
Reaction is reversible as the H–I bond formed is easily decomposed.</td></tr>
<tr><td colspan="2"></td><td>HF</td><td>HCl</td><td>HBr</td><td>HI</td></tr>
<tr><td colspan="2">Bond Energy H−X / kJ mol^{-1}</td><td>+562</td><td>+431</td><td>+366</td><td>+299</td></tr>
<tr><td rowspan="3">Thermal stability</td><td>Trend</td><td colspan="4">Thermal stability of H–X <u>decreases</u> down the group. →
Reason:
⇒ Valence orbital used in bonding is bigger and more diffused going down the group
⇒ Overlap of orbitals between the H atom and halogen atom is less effective going down the group of HF > HCl > HBr > HI
⇒ <u>H–X bond strength decreases</u></td></tr>
<tr><td>Observation</td><td colspan="2">Stable to heat ∴ do not decompose</td><td>Decomposes <u>slightly</u> on heating, <u>brown fumes of Br_2 observed</u></td><td><u>Easily</u> decomposed, gives out <u>dense purple fumes of I_2 on gentle heating</u></td></tr>
<tr><td>equation</td><td colspan="2">Nil</td><td>$2HBr\ (g) \rightarrow H_2\ (g) + Br_2\ (g)$</td><td>$2\ HI\ (g) \rightarrow H_2\ (g) + I_2\ (g)$</td></tr>
</table>

8 USES OF GROUP 17 ELEMENTS AND COMPOUNDS (SDL)

Element	Compounds and their uses
fluorine	▪ chlorofluorocarbon (CFC) compounds are used in fire extinguishers, refrigerants and aerosol propellants disadvantage: depletion of ozone layer due to free radical reactions ▪ fluorides are used in toothpaste ▪ polytetrafluoroethene, PTFE (also known as 'Teflon') used in non–stick cookware
chlorine	▪ in water purification to kill bacteria and disinfect water due to its strong and effective oxidising property ▪ manufacture of organic solvents like CCl_4 ▪ manufacture of plastics like polyvinyl chloride (PVC) ▪ manufacture of HCl ▪ manufacture of chlorates - ClO^- for bleaching clothes, paper pulp, also used as disinfectant - ClO_3^- for explosives, matches, fireworks and weed killers
bromine	▪ silver bromide for photographic film ▪ manufacture of 1,2–dibromoethane which is added to petrol together with an anti–knock agent, tetraethyl lead (IV), to remove lead as volatile lead bromide, preventing it from fouling the spark plugs
iodine	▪ preparation of dyes ▪ silver iodide in colour photography ▪ *Tincture of Iodine*, which consists of I_2 in KI diluted with 90% ethanol, is used as an antiseptic

Checkpoint 4

5

Attempt these questions and check your understanding against worked solutions on page 55.

Chemical Properties of the Group 2 elements

1 Which of the statements about Group 2 elements is correct?

A The reducing power decreases down the group.

B The ionic radius increases down the group.

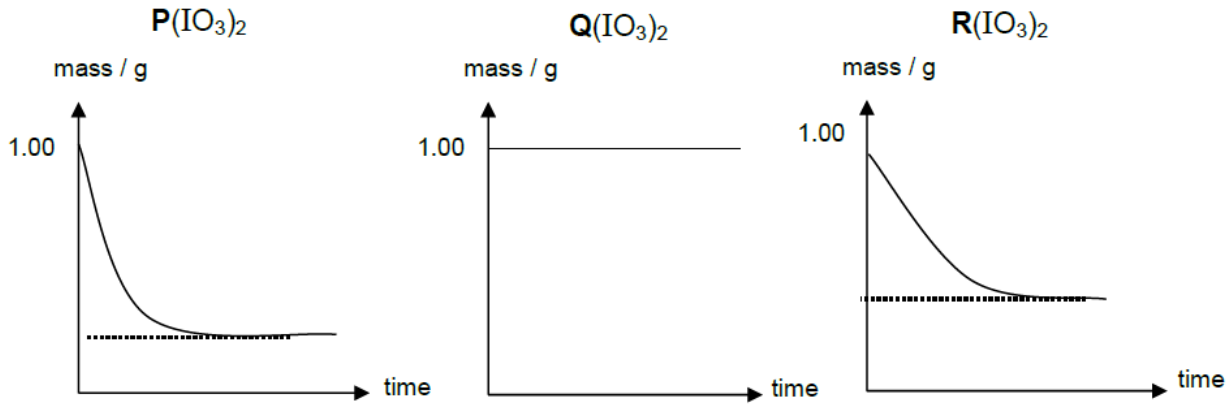
C The electronegativity increases down the group.

D The reactivity with water decreases down the group.

( )

SAJC Prelim 2007/I/14

2 **P**, **Q** and **R** are Group 2 elements. The three graphs below show the change in mass when 1.00 g each of **P**$(IO_3)_2$, **Q**$(IO_3)_2$ and **R**$(IO_3)_2$ were heated separately at a temperature of T ^oC.



Which of the following shows the elements in order of **decreasing** atomic number?

A **Q, R, P**

B **P, R, Q**

C **P, Q, R**

D **R, Q, P**

( )

ACJC Prelim 2015/I/27

Hint: Apply the concept of 'Thermal Decomposition of Group 2 Carbonates'.

Chemical Properties of the Group 17 elements

3 Which statement explains why HI has a lower thermal stability than HC*l* and HBr?

A The enthalpy change of formation of HI is the most exothermic.
B The HI bond has the lowest bond energy.
C The HI bond has the shortest length.
D The HI molecule has the lowest polarity.

( )
N2019/H1/I/17

4 The table shows the results of experiments in which the halogens X_2, Y_2 and Z_2 were added to separate aqueous solutions containing X^-, Y^- and Z^- ions.

	$X^-(aq)$	$Y^-(aq)$	$Z^-(aq)$
X_2	no reaction	no reaction	no reaction
Y_2	X_2 formed	no reaction	Z_2 formed
Z_2	X_2 formed	no reaction	no reaction

Which set contains the ions X^-, Y^- and Z^- in order of their decreasing strength as a reducing agent?

	strongest ⟶		weakest
A	X^-	Y^-	Z^-
B	X^-	Z^-	Y^-
C	Y^-	Z^-	X^-
D	Z^-	X^-	Y^-

( )
N2003/I/17

5 The halogen astatine lies below iodine in the Periodic Table. The properties of astatine can be predicted from the trends in the properties of Group 17.

Which predictions can be made for astatine?

1 It forms a hydride whose bond energy is greater than that of hydrogen iodide.
2 It is less electronegative than iodine.
3 It oxidises chloride ions to chlorine.

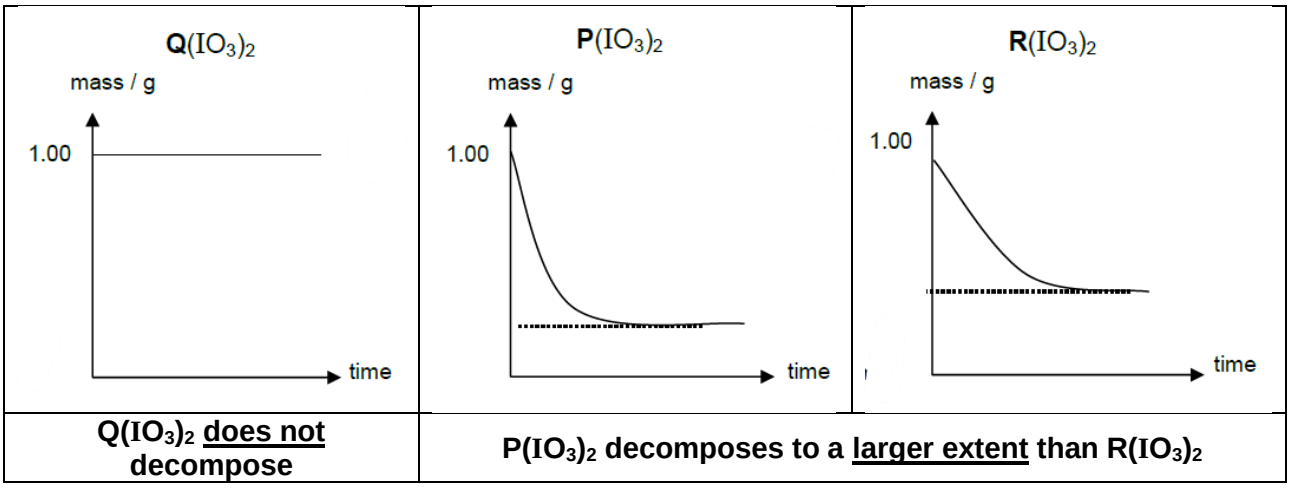
A 1 and 3
B 1 only
C 2 and 3
D 2 only

( )
N2018/H1/I/12

Solutions to Checkpoint 4

Q1: B
For group 2 metals
- Reducing power **increases** down the group
- Ionic radii **increases** down group
- Electronegativity **decreases** down the group
- Reactivity with water **increases** down the group

Q2: A
Similar to Group 2 carbonates, the **ease of thermal decomposition of Group 2 iodates, $M(IO_3)_2$ decreases** down the group.



$Q(IO_3)_2$ does not decompose	$P(IO_3)_2$ decomposes to a larger extent than $R(IO_3)_2$

∴ elements in order of decreasing (biggest to smallest) atomic number: **Q**, **R**, **P**. (No calculation needed.

Q3: B
The thermal stability of Group 17 hydrogen halides is due related to bond energy.
Explanation:

- Down the group of HC*l*, HBr and HI
- **Valence orbital** used in bonding is **bigger** and **more** diffused
- **Overlap of orbitals** between the H atom and the halogen atom is **less** effective
- **Bond energy**: HC*l* > HBr > HI
- **Thermal stability**: HC*l* > HBr > > HI

Q4: B
Halides are strong reducing agents. The stronger the reducing agent, the more likely it will undergo oxidation.
X^- is the strongest reducing agent as it is oxidised when it reacted with X_2 and Y_2.
Y^- is the weakest reducing agent since it has 'no reaction' with all halogens which suggest that it does not undergo oxidation readily.

Q5: D
Down Group 17,
1: H–X bond strength decreases (can refer to *Data Booklet*)
2: electronegativity decreases
3: Order of oxidising power of the halogens decreases
F_2 > Cl_2 > Br_2 > I_2 > At_2
Hence, At_2 is unable to displace/ oxidised Cl^- ion.