eunoia
JUNIOR COLLEGE

2022 JC1 H2 CHEMISTRY (9729)
CORE IDEA 2 − STRUCTURES AND PROPERTIES
Topic 11: THE PERIODIC TABLE

Name: ____ Civics Group: _________

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) candidates should be able to:

(a) recognise 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 ionisation 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) candidates should be able to:

(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

For elements in Group 2 (magnesium to barium) and Group 17 (chlorine to iodine) candidates should be able to:

(f) describe and deduce from $E^\ominus$ values the relative reactivity of elements of:
  (i) Group 2 as reducing agents
  (ii) Group 17 as oxidising agents
(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, candidates 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

LECTURE CONTENT

REFERENCES

1 **Chemistry (for CIE AS & A Level)** by Peter Cann & Peter Hughes

1 Introduction (INDEPENDENT STUDY)

The Periodic Table is an arrangement of elements in order of **increasing atomic (proton) number.** The properties of the elements are a periodic function of their atomic numbers (*i.e.* repeating after a regular interval).

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	s block		***d block***										***p block***					
Period																		
1								H										He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg	← **Transition metals** →										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	lanthanoids	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	actinoids	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn		Fl		Lv		

Block	Electronic Configuration
***s* block**	ns^1 to ns^2
***p* block**	ns^2np^1 to ns^2np^6
***d* block**	[Ar] $3d^{1}4s^{2}$ to [Ar] $3d^{10}4s^{2}$ [Kr] $4d^{1}5s^{2}$ to [Kr] $4d^{10}5s^{2}$ [Xe] $4f^{14}5d^{2}6s^{2}$ to [Xe] $4f^{14}5d^{10}6s^{2}$ [Rn] $5f^{14}5d^{2}6s^{2}$ to [Rn] $5f^{14}5d^{10}6s^{2}$
***f* block**	[Xe] $5d^{1}6s^{2}$ to [Xe] $4f^{14}5d^{1}6s^{2}$ [Rn] $6d^{1}7s^{2}$ to [Rn] $5f^{14}6d^{1}7s^{2}$

Note: electronic configurations stated for the s and p block are *valence* electronic configurations.

	***f* block**														
lanthanoids	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
actinoids	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Group

- Refers to the **vertical columns** of elements in the Periodic Table.
- Elements within the same Group have the **same outer electronic configuration** and hence **similar chemical properties**.

Period

- Refers to the **horizontal rows** of elements in the Periodic Table.
- Period number = Number of electron shells (or Principal Quantum Shells)
- Elements within the same Period have the **same number of electron shells**.

Periodicity

- Refers to the **recurrence of similar properties** at regular intervals when elements are arranged in increasing atomic number.

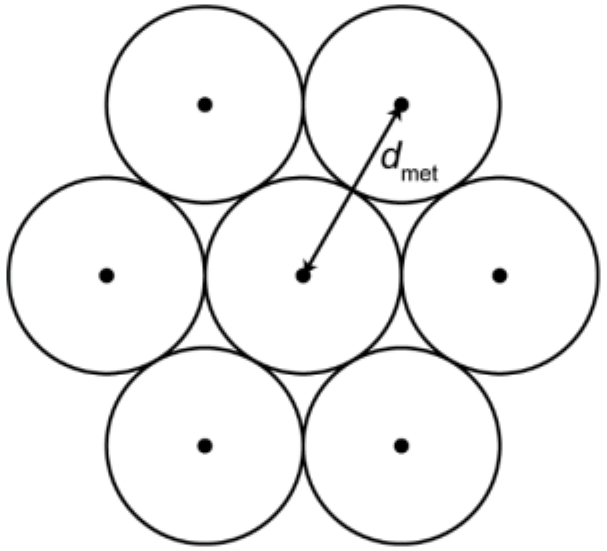
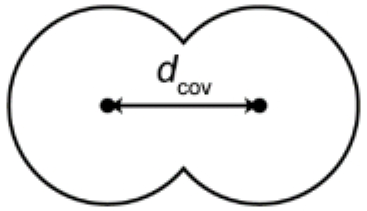
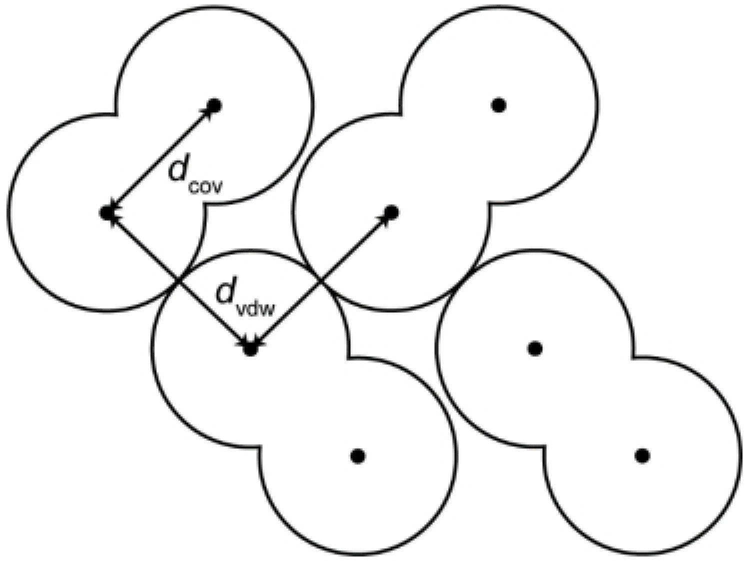
2 Trend and Properties of Period 3 Elements (INDEPENDENT STUDY)

LO (a) recognise 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 ionisation 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)

2.1 Atomic Radius

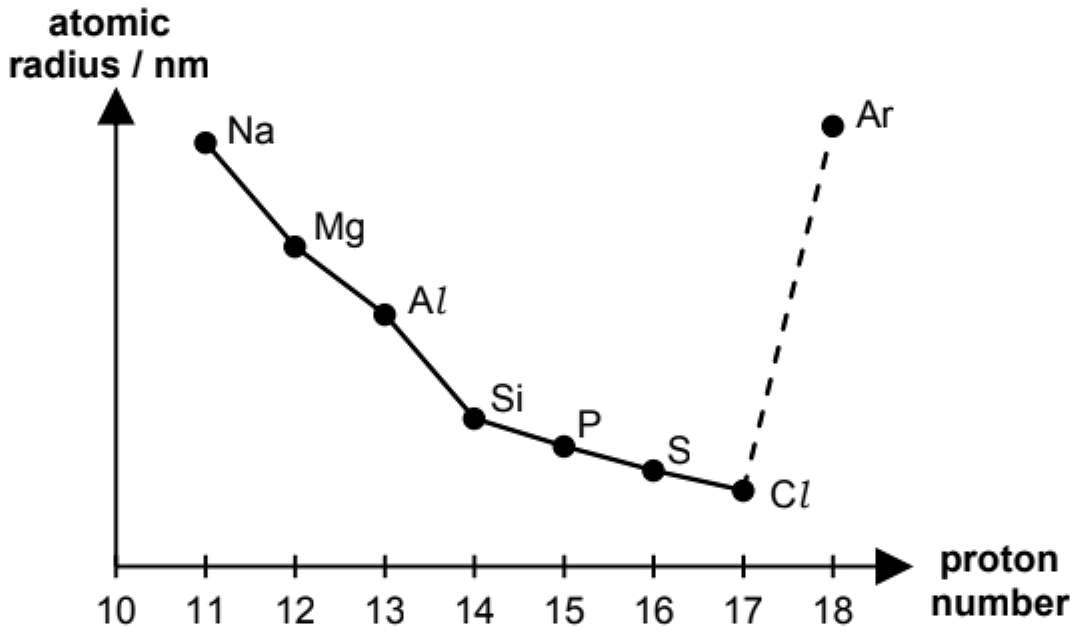
Atomic radius is defined as *half the distance* between the centres of two adjacent atoms found in the structure of the element.

It can be measured differently, depending on the inter-atomic bonding:

The metallic radius of an atom is half the distance between the **centres of two adjacent atoms in the metal**  $r_{met} = \frac{d_{met}}{2}$	The covalent radius of an atom is half the distance between the **centres of two adjacent atoms that are covalently bonded** *e.g.* within Cl_2 molecules  $r_{cov} = \frac{d_{cov}}{2}$	The van der Waals' radius of an atom is half the distance between the **centres of two adjacent atoms which are not chemically bonded** *e.g.* between Cl_2 molecules in solid Cl_2  $r_{vdw} = \frac{d_{vdw}}{2}$

For example, for Cl_2 molecules, we can consider its covalent radius (0.102 nm) and van der Waals' radius (0.175 nm). For noble gases such as argon, however, as it does not form any compounds, we will only consider its van der Waals' radius (0.190 nm). **To make meaningful comparisons between the sizes of different atoms, we will usually compare the radii of the same type**.

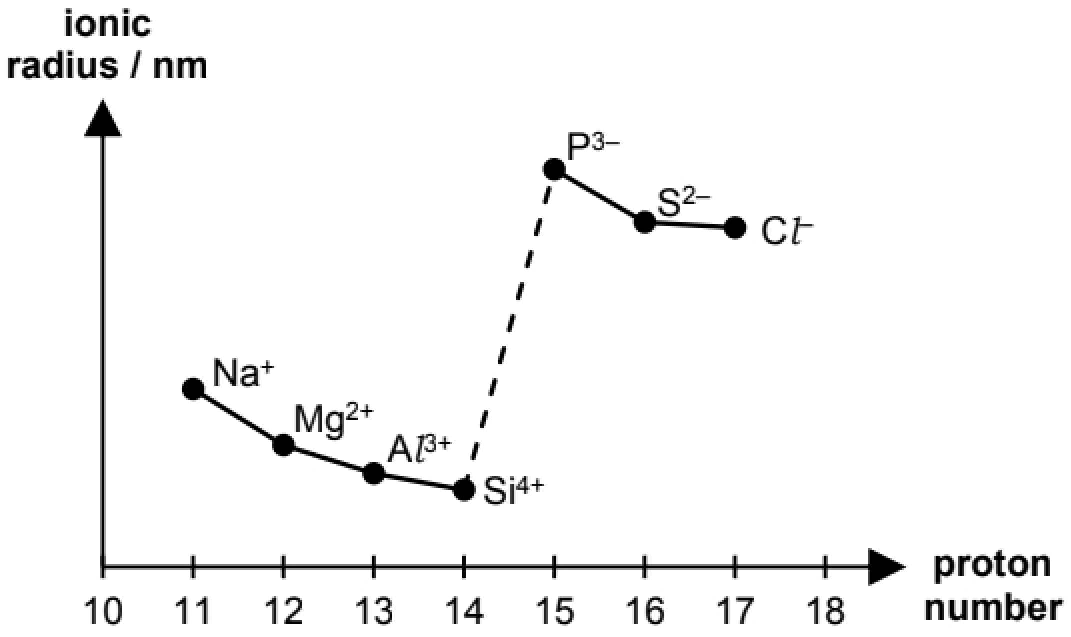
Graph 1: Trend of Atomic Radii Across Period 3



Period 3 element	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic No.	11	12	13	14	15	16	17	18
Electronic Configuration	[Ne] $3s^1$	[Ne] $3s^2$	[Ne] $3s^2\,3p^1$	[Ne] $3s^2\,3p^2$	[Ne] $3s^2\,3p^3$	[Ne] $3s^2\,3p^4$	[Ne] $3s^2\,3p^5$	[Ne] $3s^2\,3p^6$
Atomic Radius	Across the period, **atomic radii decrease** because: ❋ Number of protons in the nucleus of the atom increases, thus **nuclear charge increases** ❋ Same number of inner shell electrons, thus **shielding effect is approximately constant** ❋ **Increase in effective nuclear charge** ❋ **Increase in electrostatic forces of attraction** between the nucleus and valence electron							❋ r_{Ar} appears to be the largest as value (0.190 nm) in *Data Booklet* is the **van der Waals' radius**.

2.2 Ionic Radius

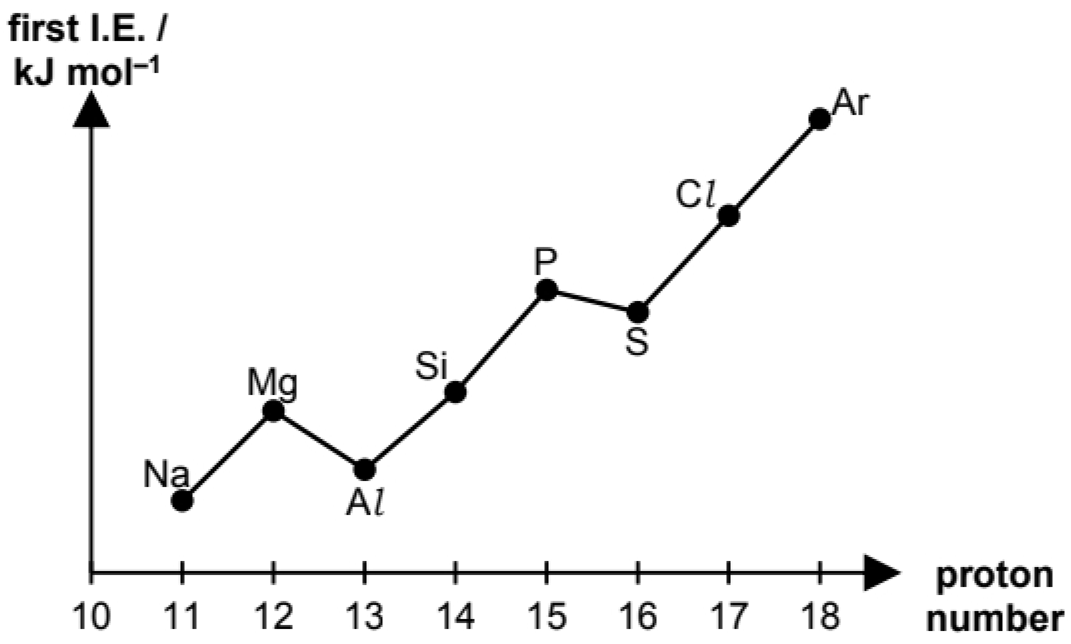
Graph 2: Trend of Ionic Radii Across Period 3



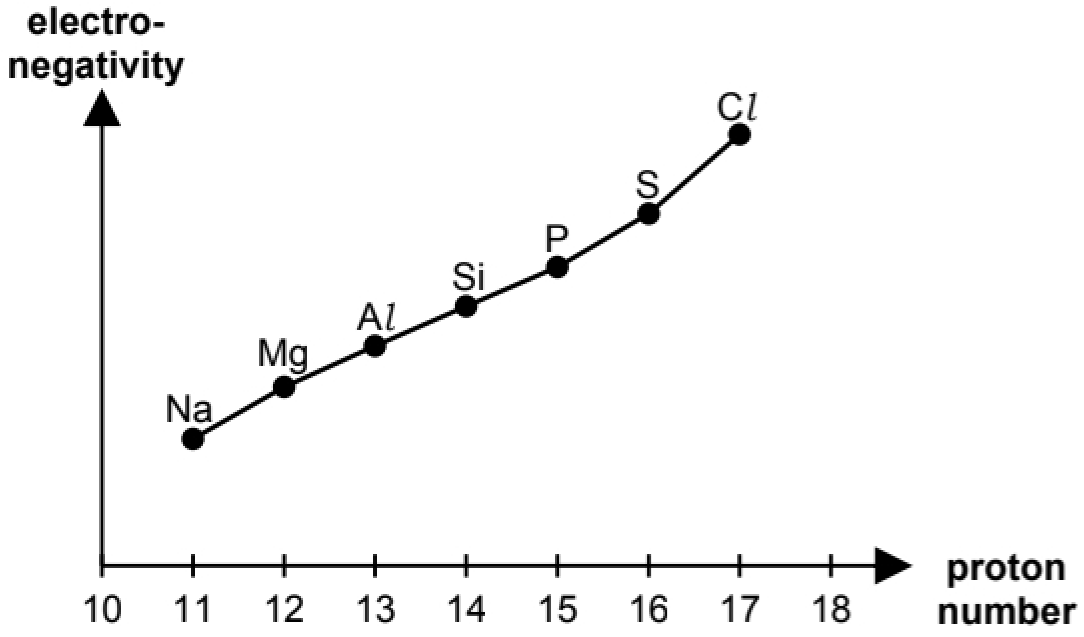
Period 3 element	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic No.	11	12	13	14	15	16	17	18
Period 3 ions	$_{11}Na^{+}$	$_{12}Mg^{2+}$	$_{13}Al^{3+}$	$_{14}Si^{4+}$	$_{15}P^{3-}$	$_{16}S^{2-}$	$_{17}Cl^{-}$	
Electronic Configuration	$[He]\ 2s^2\ 2p^6$	$[He]\ 2s^2\ 2p^6$	$[He]\ 2s^2\ 2p^6$	$[He]\ 2s^2\ 2p^6$	$[Ne]\ 3s^2\ 3p^6$	$[Ne]\ 3s^2\ 3p^6$	$[Ne]\ 3s^2\ 3p^6$	
Ionic Radius	From Na^{+} to Si^{4+}, **ionic radii decrease** because: ⚛ Nuclear charge **increases** ⚛ **Same** inter-electronic repulsion since the cations are isoelectronic (same number of electrons) ⚛ **Stronger** electrostatic forces of attraction between the nucleus and the valence electrons				From P^{3-} to Cl^{-}, **ionic radii decrease** because: ⚛ Nuclear charge **increases** ⚛ **Same** inter-electronic repulsion since the anions are isoelectronic (same number of electrons) ⚛ **Stronger** electrostatic forces of attraction between the nucleus and the valence electrons				
	Size of anions > cations because: ⚛ Anions have **one more filled principal quantum shell** compared to cations ⚛ Valence electrons in anions are **further** from the nucleus than those in cations ⚛ **Weaker** electrostatic forces of attraction between the nucleus and valence electrons in anions than in cations							

2.3 1st Ionisation Energy and Electronegativity

Graph 3: Trend of 1st Ionisation Energies Across Period 3



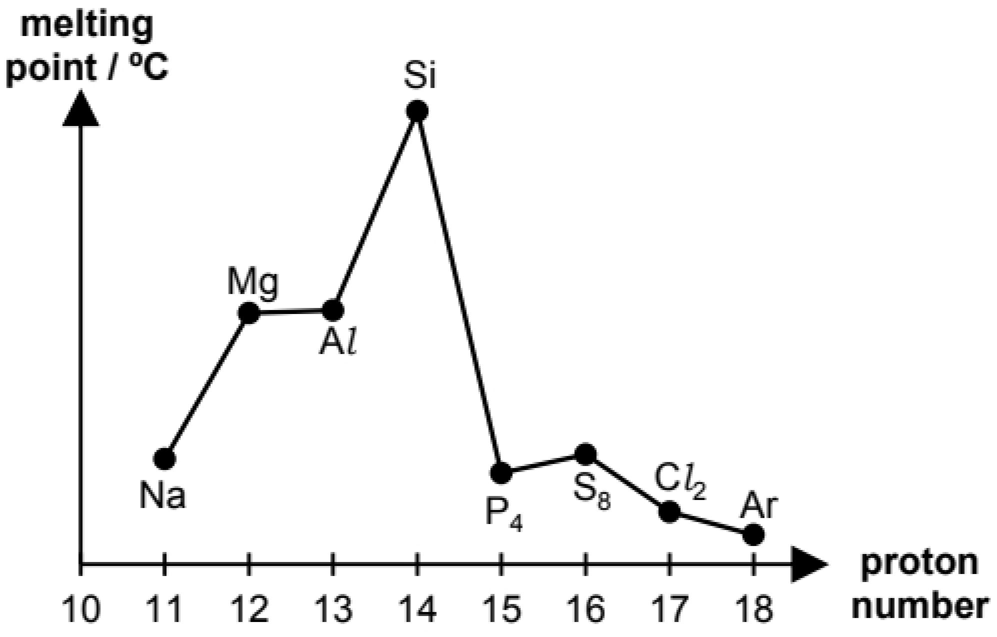
Graph 4: Trend of Electronegativities Across Period 3



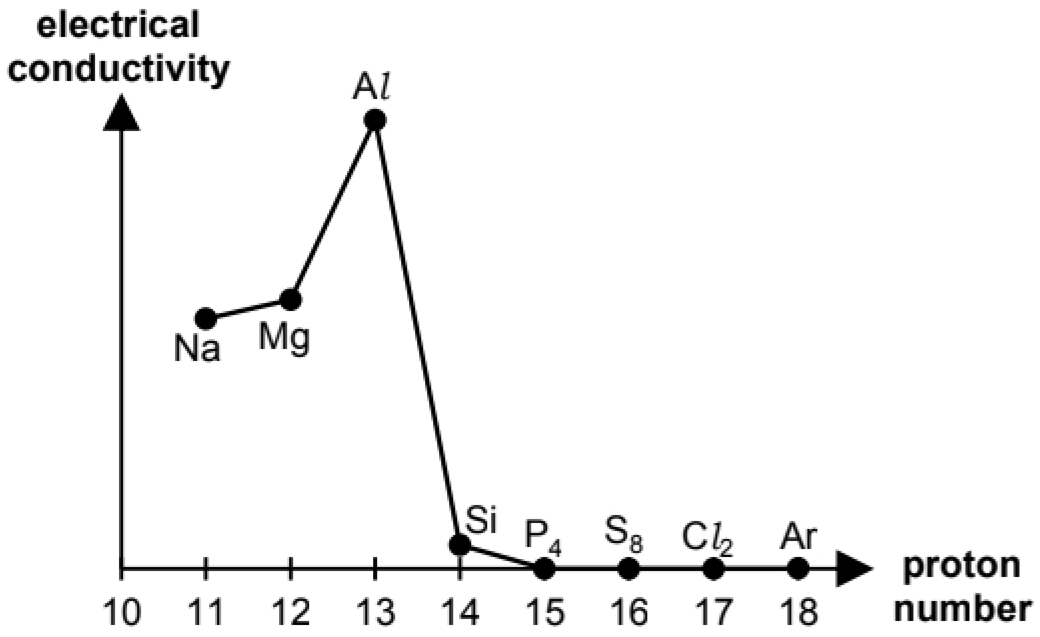
<table>
<tr><th>Period 3 element</th><th>Na</th><th>Mg</th><th>Al</th><th>Si</th><th>P</th><th>S</th><th>Cl</th><th>Ar</th></tr>
<tr><td>Atomic No.</td><td>11</td><td>12</td><td>13</td><td>14</td><td>15</td><td>16</td><td>17</td><td>18</td></tr>
<tr><td>Electronic Configuration</td><td>[Ne] $3s^1$</td><td>[Ne] $3s^2$</td><td>[Ne] $3s^2 3p^1$</td><td>[Ne] $3s^2 3p^2$</td><td>[Ne] $3s^2 3p^3$</td><td>[Ne] $3s^2 3p^4$</td><td>[Ne] $3s^2 3p^5$</td><td>[Ne] $3s^2 3p^6$</td></tr>
<tr><td>1st Ionisation Energy</td><td colspan="8">Across the period, 1st IE generally increase because:
⚛ Number of protons in the nucleus of the atom increases, thus nuclear charge increases
⚛ Same number of inner shell electrons, thus shielding effect is approximately constant
⚛ Increase in effective nuclear charge
⚛ Increase in electrostatic forces of attraction between the nucleus and valence electron
⚛ More energy needed to remove the valence electron, increase in 1st I.E. across the period
1st IE of Al is lower than Mg because:
⚛ The 3p electron to be removed from Al is at a higher energy level than the 3s electron to be removed from Mg.
⚛ Less energy is required to remove the 3p electron from Al.
1st IE of S is lower than P because:
⚛ There is inter-electronic repulsion between the paired electrons in the same orbital for S.
⚛ Less energy is thus required to remove a 3p electron from S due to inter-electron repulsion between the paired electrons in the same orbital.</td></tr>
<tr><td>Electronegativity</td><td colspan="8">Across the period, electronegativity increases because:
⚛ Number of protons in the nucleus of the atom increases, thus nuclear charge increases
⚛ Same number of inner shell electrons, thus shielding effect is approximately constant, increase in effective nuclear charge
⚛ Increase in electrostatic forces of attraction between the nucleus and valence electron</td></tr>
</table>

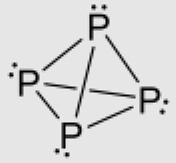
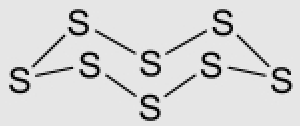
2.4 Melting Point and Electrical Conductivity

Graph 5: Trend of Melting Points Across Period 3



Graph 6: Trend of Electrical Conductivities Across Period 3



Period 3 element	**Na**	**Mg**	**Al**	**Si**	**P** (P_4) 	**S** (S_8) 	**Cl** (Cl_2) Cl—Cl	**Ar**
Structure	Giant metallic lattice structure			Giant molecular	Simple molecular structure			
Melting Point	**High melting points** because: • **Large** amount of energy is needed to overcome the strong **metallic bonds**. **Melting points of Al > Mg > Na** because: • Increase in number of **delocalised valence electrons** contributed per atom • **Smaller cationic radius** • **Higher charge density** $\left(\frac{q^+}{r_+}\right)$ • **More energy is needed to overcome** the **strong electrostatic forces of attraction** between **metal cations and sea of delocalised electrons**			**High melting point** because: • **Large** amount of energy is needed to overcome the **strong and extensive covalent bonds** between atoms	**Low melting points** because: • **Small** amount of energy is needed to overcome the **weak intermolecular instantaneous dipole–induced dipole interactions** **Melting points of S_8 > P_4 > Cl_2 > Ar** because: • **Increase** in **number of electrons in the molecule** • **Increase** in **polarisability of electron cloud** • **More** energy required to overcome **stronger intermolecular instantaneous dipole-induced dipole interactions**			
Electrical Conductivity	**High electrical conductivity** because: • Presence of sea of **delocalised electrons** as **mobile** charge carriers to conduct electricity.			**Low electrical conductivity** because: • Si is a metalloid.	**Non-conductors** of electricity in any state because: • Electrons are **localised** in **covalent bonds** and **not mobile** to conduct electricity.			

3 Trend and Properties of Period 3 Oxides and Chlorides

3.1 Oxidation numbers

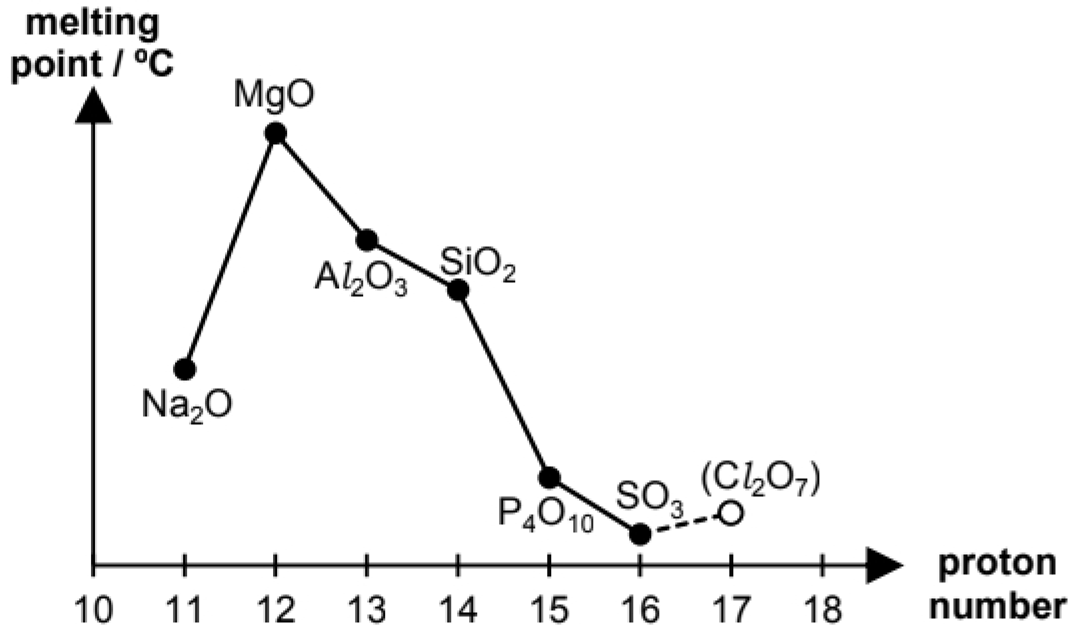
LO (d) (i) state and explain the variation in the highest oxidation number of the elements in oxides (for Na_2O; MgO; Al_2O_3; SiO_2; P_4O_{10}; SO_3) and chlorides (for $NaCl$; $MgCl_2$; $AlCl_3$; $SiCl_4$; PCl_5)

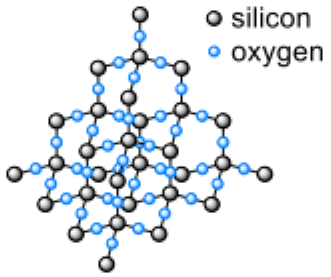
Period 3 element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**
Oxide	Na_2O	MgO	Al_2O_3	SiO_2	(P_4O_6) P_4O_{10}	(SO_2) SO_3
Chloride	$NaCl$	$MgCl_2$	$AlCl_3$	$SiCl_4$	(PCl_3) PCl_5	
Oxidation number	+1	+2	+3	+4	(+3) +5	(+4) +6
	Across the period, **oxidation number increases** because: ❈ Number of valence electrons available for bond formation increases. ❈ Maximum oxidation number = number of valence electrons of an atom **P and S can exhibit multiple oxidation numbers** because: ❈ Presence of **vacant and energetically accessible d orbitals** which can be used for bonding through **expansion of the octet structure**.					

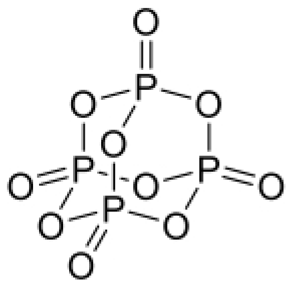
3.2 Melting Points

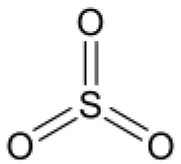
LO	(d)	(i)	state and explain the variation in bonding in oxides and chlorides in terms of electronegativity (with the exception of $AlCl_3$)
		(vi)	suggest the types of structure and bonding present in the oxides and chlorides from observations of their chemical and physical properties

Graph 7: Trend of Melting Points of Period 3 Oxides



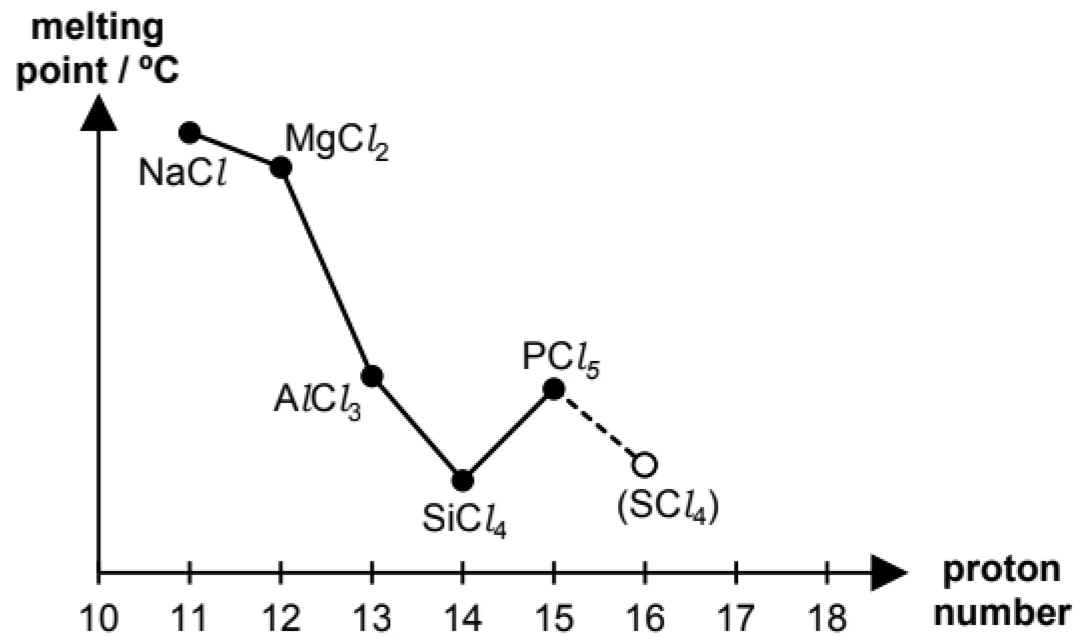






Period 3 oxide	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_4O_{10}**	**SO_3**
Structure	Giant ionic structure			Giant molecular	Simple molecular structure	
Melting point of oxides	**High melting points** because: - **Large** amount of energy is needed to overcome the **strong ionic bonds**. **Melting points of MgO > Na_2O** because: - $\lvert\Delta H_{latt}\rvert \propto \left\lvert\frac{q^+ \times q^-}{r_+ + r_-}\right\rvert$ - Anionic charge and anionic radius remain the same, but Mg^{2+} has **higher ionic charge and smaller ionic radius** than Na^+. - **More** energy is required to overcome the **stronger ionic bonds** in MgO. **Melting points of MgO > Al_2O_3** because: - Al^{3+} has **higher ionic charge and smaller ionic radius** than Mg^{2+} - Al^{3+} has a **higher charge density** and hence **higher polarising power** - Al_2O_3 has some **covalent character**, **weakening** the **ionic bonds** in Al_2O_3.			**High melting point** because: - **Large** amount of energy is needed to overcome the **strong and extensive covalent bonds** between atoms	**Low melting points** because: - **Small** amount of energy is needed to overcome the **weak intermolecular instantaneous dipole–induced dipole interactions** **Melting points of P_4O_{10} > SO_3** because: - **More electrons** in P_4O_{10} molecule than SO_3 - **Increase** in **polarisability of electron cloud** - **More** energy required to overcome **stronger intermolecular instantaneous dipole-induced dipole interactions**	

Graph 8: Trend of Melting Points of Period 3 Chlorides

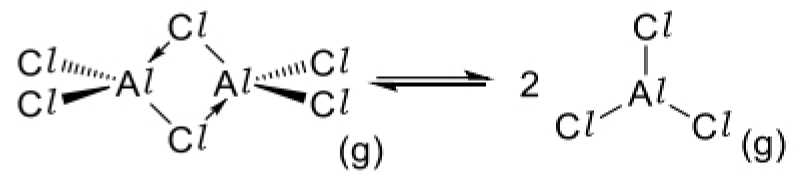


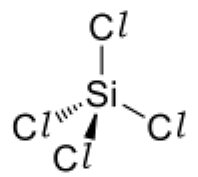
FYI: More information about AlCl_3

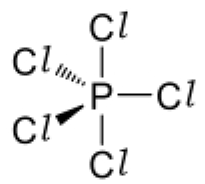
- At temperatures above 180 °C, the chloride exists as a molecular **dimer, Al_2Cl_6** held together by **instantaneous dipole-induced dipole attraction**.
- The dimer persists in the gaseous phase when the liquid boils:

$$Al_2Cl_6\,(l) \xrightleftharpoons{\text{heat}} Al_2Cl_6\,(g)$$

- At higher temperatures, **gaseous Al_2Cl_6 dissociates to give monomeric $AlCl_3$ molecules**:





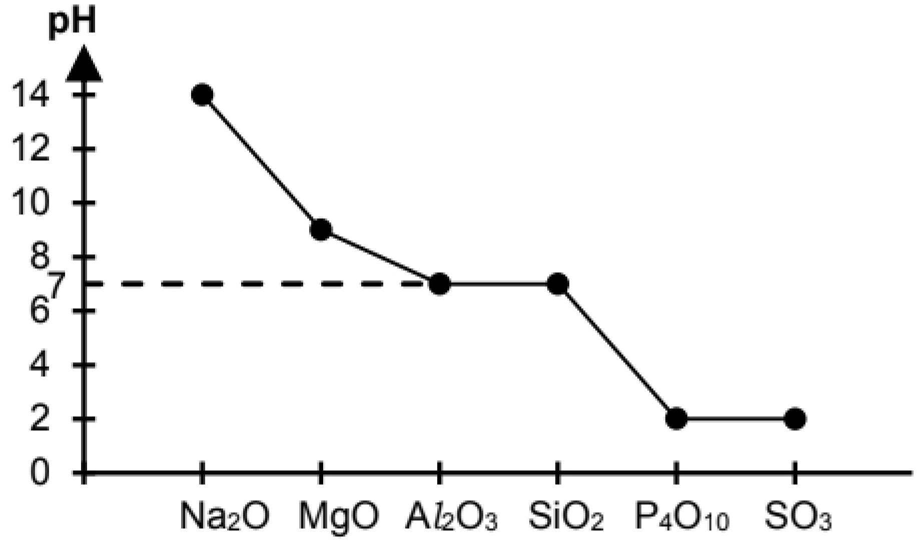


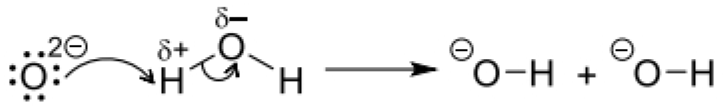
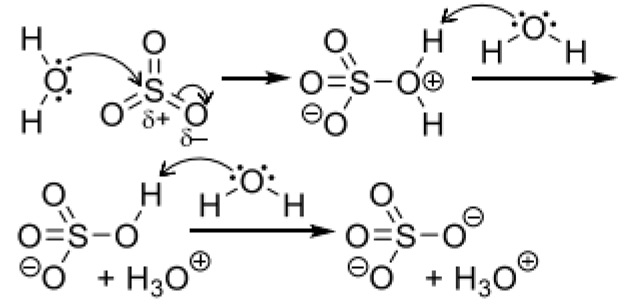
Period 3 chloride	NaCl	MgCl_2	AlCl_3	SiCl_4	PCl_5
Structure	Giant ionic structure		Simple molecular structure		
			- Al^{3+} has **high ionic charge and small ionic radius** - Al^{3+} has a **high charge density** and hence **high polarising power** - Al–Cl bond has **significant covalent character**, hence AlCl_3 has a **simple molecular structure**		
Melting point of chlorides	**High melting points** because: - **Large** amount of energy is needed to overcome the **strong ionic bonds**. **Melting points of NaCl > MgCl_2** because: - Mg^{2+} has **higher ionic charge and smaller ionic radius** than Na^+ - Mg^{2+} has a **higher charge density** and hence **higher polarising power** - MgCl_2 has some **covalent character**, **weakening** the **ionic bonds** in MgCl_2.		**Low melting points** because: - **Small** amount of energy is needed to overcome the **weak intermolecular instantaneous dipole–induced dipole interactions** **Melting points of PCl_5 > SiCl_4** because: - **More electrons** in PCl_5 molecule than SiCl_4 - **Increase** in **polarisability of electron cloud** - **More** energy required to overcome **stronger intermolecular instantaneous dipole-induced dipole interactions**		

3.3 Reactions of Period 3 Oxides with Water, Acids and NaOH

LO (d)	(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

Graph 9: Trend of pH of Resulting Solution when Period 3 Oxides React with Water



Period 3 oxide	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_4O_{10}**	**SO_3**
Structure	Giant ionic structure			Giant molecular structure	Simple molecular structure	
Reaction with water	**<u>Dissolves completely</u>** in water to form NaOH $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$	**<u>Dissolves partially in water</u>** to form $Mg(OH)_2$ due to ***high lattice energy*** of MgO $MgO(s) + H_2O(l) \rightleftharpoons Mg(OH)_2(s)$	**<u>No reaction</u>** **Al_2O_3 <u>does not dissolve</u>** in water because of its extremely **<u>high lattice energy</u>. Large amount** of energy is required to break the **strong ionic bonds**.	**<u>No reaction</u>** **SiO_2 <u>does not dissolve</u>** in water because **large amount** of energy is required to break the **strong and extensive covalent bonds** between **Si and O atoms**.	**<u>Dissolves</u>** in water to form H_3PO_4 $P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$	**<u>Dissolves</u>** in water to form H_2SO_4 $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$
	Due to presence of the basic O^{2-} anion, which reacts with H_2O by accepting a H^+ to give OH^-. **FYI: Mechanism of reaction between basic oxide and water**  $(O^{2-} + H_2O \rightarrow 2OH^-)$ $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$ ($Mg(OH)_2(s)$ is sparingly soluble)				Due to presence of $\overset{\delta+}{X}=\overset{\delta-}{O}$ (X = S, P), which reacts with H_2O to give X—O—H bonds, which can dissociate to give H^+. **FYI: Mechanism of reaction between an acidic oxide, SO_3 and water**  	
	pH 14	pH 9	pH 7	pH 7	pH 2	pH 2

Period 3 oxide	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_4O_{10}**	**SO_3**
Structure	Giant ionic structure			Giant molecular structure	Simple molecular structure	
Nature of oxide	Basic		Amphoteric (*i.e.* shows both acidic and basic properties)	Acidic		
Reaction with acid	Forms **salt and water** $Na_2O(s) + 2HCl(aq)$ $\rightarrow 2NaCl(aq) + H_2O(l)$	Forms **salt and water** $MgO(s) + 2HCl(aq)$ $\rightarrow MgCl_2(aq) + H_2O(l)$	Forms **salt and water** $Al_2O_3(s) + 6HCl(aq)$ $\rightarrow 2AlCl_3(aq) + 3H_2O(l)$	**No reaction**		
Reaction with base	**No reaction**		Dissolves in **excess NaOH** to form **colourless complex, $Al(OH)_4^-$ (aq)** $Al_2O_3(s) + 2NaOH(aq)$ $+ 3H_2O(l)$ $\rightarrow 2NaAl(OH)_4$ (aq)	Reacts with **hot conc. NaOH** to form **$Na_2SiO_3(aq)$** $SiO_2(s) + 2NaOH(conc)$ $\rightarrow Na_2SiO_3(aq)$ $+ H_2O(l)$	**Forms** salt and water $P_4O_{10}(s) + 12NaOH(aq)$ $\rightarrow 4Na_3PO_4(aq) +$ $6H_2O(l)$	Forms **salt and water** $SO_3(g) + 2NaOH(aq)$ $\rightarrow Na_2SO_4(aq) + H_2O(l)$

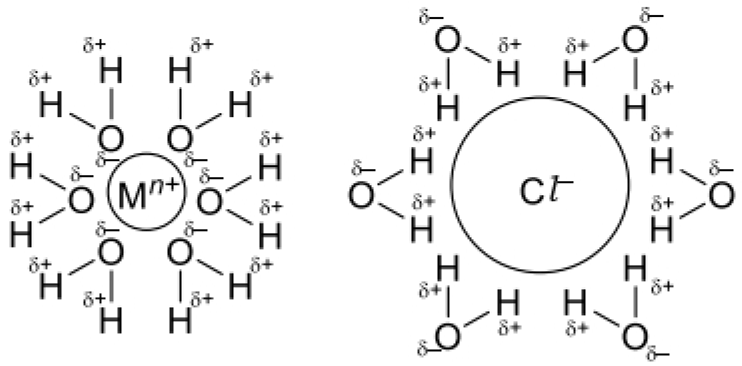
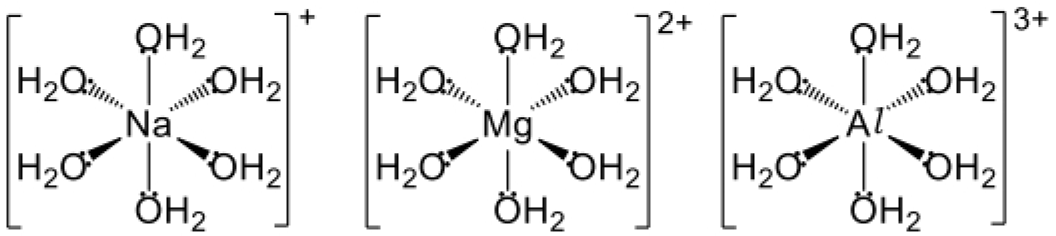
Period 3 hydroxide	**$NaOH$**	**$Mg(OH)_2$**	**$Al(OH)_3$**
Structure	Giant ionic structure		
Nature of oxide	Basic		Amphoteric (*i.e.* shows both acidic and basic properties)
Reaction with acid	Forms **salt and water** $NaOH(aq) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l)$	Forms **salt and water** $Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$	Forms **salt and water** $Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$
Reaction with base	**No reaction**		Dissolves in **excess NaOH** to form **colourless complex, $Al(OH)_4^-$ (aq)** $Al(OH)_3(s) + OH^-(aq) \rightarrow Al(OH)_4^-(aq)$

3.4 Reactions of Period 3 Chlorides with Water

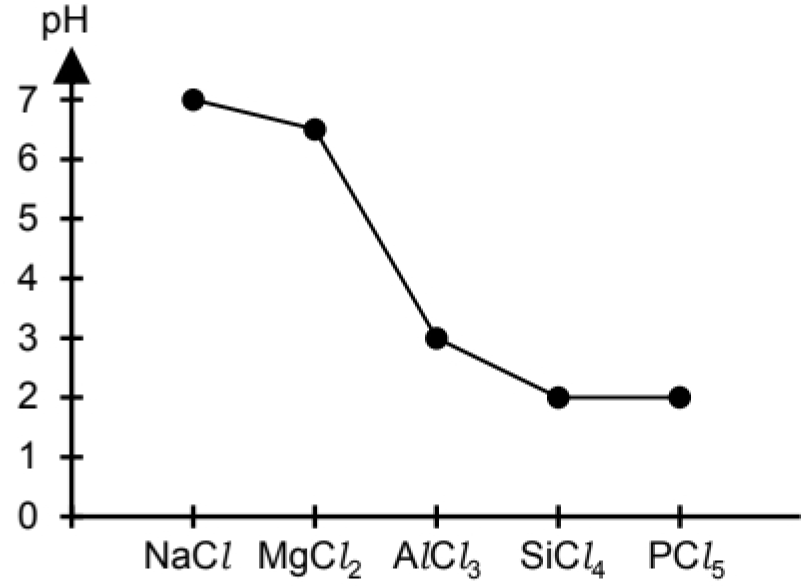
LO (d) (v) describe and explain the reactions of the chlorides with water (for $NaCl$; $MgCl_2$; $AlCl_3$; $SiCl_4$; PCl_5)

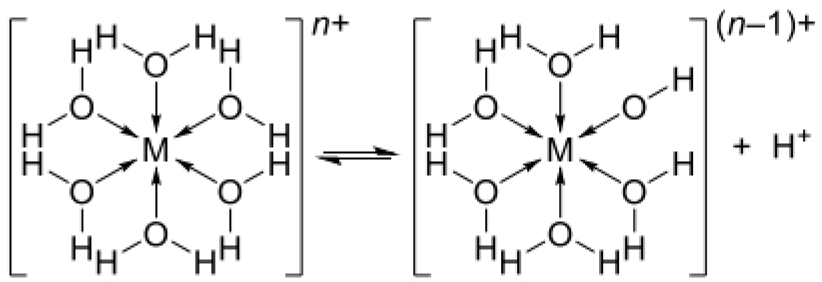
The reactions of Period 3 chlorides with water involve **hydration** and **hydrolysis** to different extents.

Across the third period **as the chlorides become more covalent, tendency to undergo hydrolysis increases**.

<table>
<tr><th>Period 3 chlorides</th><th>$NaCl$</th><th>$MgCl_2$</th><th colspan="2">$AlCl_3$</th><th>$SiCl_4$</th><th>PCl_5</th></tr>
<tr><td>Structure</td><td colspan="3">Ionic</td><td colspan="3">Simple molecular</td></tr>
<tr><td>Hydration
The physical attraction between the cations and anions with the negative dipole and positive dipole of the water molecules, respectively
(<i>i.e.</i> the formation of ion-dipole interactions).</td><td colspan="3">Undergoes hydration to give the hydrated metal ions and chloride ions
• $MCl_n(s) \xrightarrow{H_2O(l)} M^{n+}(aq) + nCl^-(aq)$

• The hydrated metal ions $Na^+(aq)$, $Mg^{2+}(aq)$ and $Al^{3+}(aq)$ are complexes where water molecules surround and form dative bonds with the cation:

The metal ion uses one empty 3s, three 3p and two 3d orbitals to form six hybrid orbitals, that can accept the six lone pairs from 6 H_2O molecules.</td><td colspan="3">Does not undergo hydration</td></tr>
<tr><td rowspan="2">Hydrolysis
The chemical reaction of the cation or covalent chlorides and water molecules such that the O–H bond in the water molecule is cleaved, giving H_3O^+ ions.</td><td>$[Na(H_2O)_6]^+$ does not undergo hydrolysis</td><td>$[Mg(H_2O)_6]^{2+}$ undergoes slight hydrolysis</td><td>$[Al(H_2O)_6]^{3+}$ undergoes extensive hydrolysis</td><td rowspan="2" colspan="3">Undergoes complete hydrolysis
• Si and P atoms have vacant and energetically accessible 3d orbitals
• Al atom have vacant and energetically accessible 3p orbital
to accept the lone pair from H_2O molecules, forming a dative bond.
<i>Think:</i> Does CCl_4 undergo hydrolysis?
CCl_4 does not hydrolyse in water as C has no energetically accessible vacant orbitals for dative bonding with H_2O molecules.
In addition, C atom is small and is shielded by larger Cl atoms. H_2O nucleophile could not easily attack the central C atom due to steric effect.</td></tr>
<tr><td></td><td colspan="2">• Relatively higher charge density of metal cation polarises the electron cloud of the surrounding water molecules, weakening and breaking the O–H bond, resulting in the release of ~~a~~ proton.
• $[Al(H_2O)_6]^{3+}$ undergoes more extensive hydrolysis than $[Mg(H_2O)_6]^{2+}$ as Al^{3+} has a higher charge density than Mg^{2+}.</td></tr>
</table>

Graph 10: Trend of pH of Resulting Solution when Period 3 Chlorides React with Water



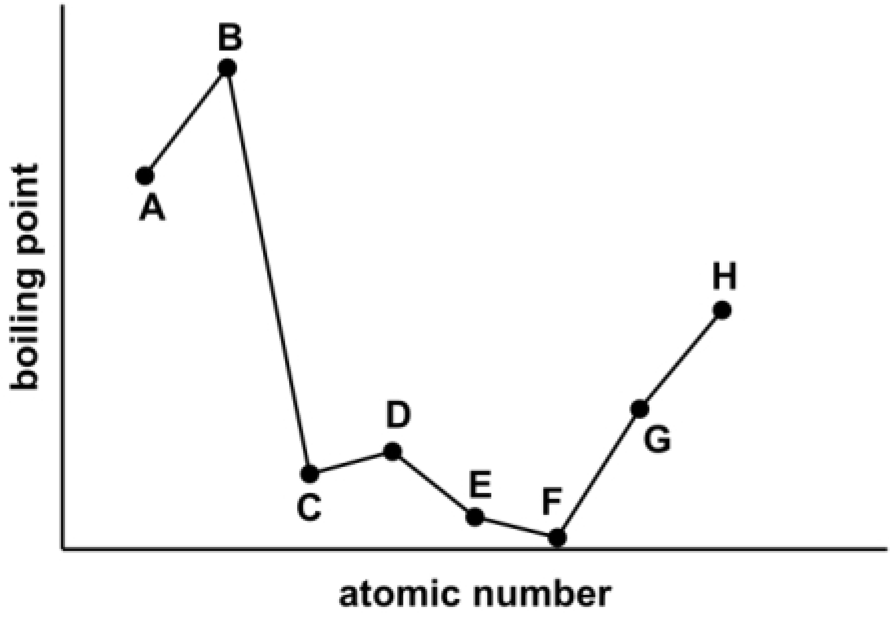
Period 3 chlorides		**NaCl**	**MgCl₂**	**AlCl₃**		**SiCl₄**	**PCl₅**
Observations		**Dissolves completely** in water	**Dissolves completely** in water	**Dissolves completely** in **excess** water	**Reacts vigorously** with water, and forms **white solid** with **white fumes of HCl** in **limited** amount of water	**Reacts vigorously** with water, and forms **white solid** with **white fumes of HCl**	**Reacts vigorously** with water, and forms **white fumes of HCl**
Reactions	**hydration**	$NaCl(s) + aq \rightarrow Na^+(aq) + Cl^-(aq)$ *(hydration)*	$MgCl_2(s) + 6H_2O(l) \rightarrow [Mg(H_2O)_6]^{2+}(aq) + 2Cl^-(aq)$ *(hydration)*	**In excess water:** $AlCl_3(s) + 6H_2O(l) \rightarrow [Al(H_2O)_6]^{3+}(aq) + 3Cl^-(aq)$ *(hydration)*			
	hydrolysis	– *(no hydrolysis)*	$[Mg(H_2O)_6]^{2+}(aq) \rightleftharpoons [Mg(H_2O)_5(OH)]^+(aq) + H^+(aq)$ *(slight hydrolysis)* 	$[Al(H_2O)_6]^{3+}(aq) \rightleftharpoons [Al(H_2O)_5(OH)]^{2+}(aq) + H^+(aq)$ *(extensive hydrolysis)*	**In limited supply of water:** $AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(s) + 3HCl(g)$ *(complete hydrolysis)* *or* $2AlCl_3(s) + 3H_2O(l) \rightarrow Al_2O_3(s) + 6HCl(g)$ *(complete hydrolysis)*	$SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(aq)$ *(complete hydrolysis)*	$PCl_5(s) + 4H_2O(l) \rightarrow H_3PO_4(aq) + 5HCl(aq)$ *(complete hydrolysis)*
pH of solution		pH 7	pH 6.5	pH 3		pH 2	pH 2

Checkpoints for Section 3

- To explain in terms of structure and bonding the physical properties of Period 3 elements and its oxides and chlorides.
- To 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)
- To state and explain the variation in bonding in oxides and chlorides in terms of electronegativity (with the exception of $AlCl_3$)
- To describe the reactions of the oxides with water (for Na_2O; MgO; Al_2O_3; SiO_2; P_4O_{10}; SO_3), including writing of equations
- To 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, including writing of equations
- To describe and explain the reactions of the chlorides with water (for $NaCl$; $MgCl_2$; $AlCl_3$; $SiCl_4$; PCl_5), including writing of equations
- To suggest the types of structure and bonding present in the oxides and chlorides from observations of their chemical and physical properties

Example 3A

1 The graph below shows the variation in the boiling points for eight consecutive elements in the Periodic Table, all with atomic number between 10 and 20. Which of the following statements is true?



A Element **D** forms an acidic oxide.

B Element **G** does not conduct electricity.

C Element **A** and beryllium are in the same group.

D The element immediately preceding **A** has a smaller atomic radius than **A**.

HCI2011/1/13

Example 3A (continued)

2 Describe the reactions, if any, of the chlorides $NaCl$, $AlCl_3$ and PCl_5 with water.
Write equations for all reactions that occur, and suggest the pH of the resulting solutions.
Relate the reactivity of these chlorides to their structure and bonding.

$NaCl$ has a ____________________, held together by ____. As a result, $NaCl$ ____________________, where the ions undergoes hydration:

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

Due to the very low charge density of the hydrated $Na^+(aq)$ ion, no hydrolysis takes place and the pH of the resulting solution is ___.

$AlCl_3$ has a ____________________, containing intramolecular polar Al–Cl covalent bonds. As a result, $AlCl_3$ ____________________with water.

In limited supply of water:

$AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(s) + 3HCl(g)$ or

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

It undergoes ____________________to give white solid with white fumes of HCl.

Some of the HCl dissolves in water to give a pH of ___ for the resulting solution.

In excess water:

$AlCl_3$ ____________________.

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

(hydration)

$[Al(H_2O)_6]^{3+}(aq) \rightleftharpoons [Al(H_2O)_5(OH)]^{2+}(aq) + H^+(aq)$

(extensive hydrolysis)

Due to the extensive hydrolysis of $[Al(H_2O)_6]^{3+}(aq)$, the pH of the resulting solution is . ___

PCl_5 has a ____________________, consisting of PCl_5 molecules containing intramolecular polar P–Cl covalent bonds. As a result, PCl_5 ____________________with water.

$PCl_5(s) + 4H_2O(l) \rightarrow H_3PO_4(aq) + 5HCl(aq)$

It undergoes ____________________to give white fumes of HCl.

Due to some of the HCl dissolving in the water with the H_3PO_4 formed, the resulting solution is highly acidic at pH ___.

N19/3/4a

4 Diagonal Relationship

LO (e) predict the characteristic properties of an element in a given Group by using knowledge of chemical periodicity

A diagonal relationship is said to exist between certain pairs of diagonally adjacent elements in the **second and third periods** of the Periodic Table.
Each of these pairs (Li and Mg, Be and A*l*, and B and Si) are said to exhibit similar properties and hence they have **diagonal relationships.**

Group 1	Group 2	Group 13	Group 14
Li ↘	Be ↘	B ↘	C
Na	Mg	A*l*	Si

Diagonal relationships occur because

(i) the elements have **similar <u>electronegativity</u>** (electronegativity increases across the period and decreases down the group)

(ii) their cations have **similar <u>charge density</u>** (charge increases across the period and radius increases down the group)

<u>Diagonal Relationship between Be and A*l*</u>

Be and A*l* have **similar electronegativity**, and Be^{2+} and Al^{3+} have **similar charge density**.

When $BeCl_2$ and $AlCl_3$ dissolved in water, both $Be^{2+}(aq)$ and $Al^{3+}(aq)$ aqua complex ions undergo **hydrolysis** to give an **acidic solution**.		Both BeO and Al_2O_3 are **amphoteric oxides** that react with both acids and bases.	
$BeCl_2 + 4H_2O \rightarrow [Be(H_2O)_4]^{2+} + 2Cl^-$ $[Be(H_2O)_4]^{2+} \rightleftharpoons [Be(H_2O)_3(OH)]^+ + H^+$	$AlCl_3 + 6H_2O \rightarrow [Al(H_2O)_6]^{3+} + 3Cl^-$ $[Al(H_2O)_6]^{3+} \rightleftharpoons [Al(H_2O)_5(OH)]^{2+} + H^+$	$BeO + 2HCl \rightarrow BeCl_2 + H_2O$ $BeO + 2NaOH + H_2O \rightarrow Na_2Be(OH)_4$	$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ $Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2NaAl(OH)_4$

<u>Diagonal Relationship between Li and Mg</u>

Group 1 nitrates decompose to give nitrite and oxygen: $NaNO_3 \rightarrow NaNO_2 + \frac{1}{2}O_2$ However, $LiNO_3$ is similar to $Mg(NO_3)_2$ which **decomposes on heating to give metal oxide**, NO_2 and O_2.		Group 1 metal hydroxides and metal carbonate are thermally stable. However, LiOH and Li_2CO_3 **decompose to form oxides** similar to $Mg(OH)_2$ and $MgCO_3$.	
$2LiNO_3 \rightarrow Li_2O + 2NO_2 + \frac{1}{2}O_2$	$Mg(NO_3)_2 \rightarrow MgO + 2NO_2 + \frac{1}{2}O_2$	$2LiOH \rightarrow Li_2O + H_2O$ $Li_2CO_3 \rightarrow Li_2O + CO_2$	$Mg(OH)_2 \rightarrow MgO + H_2O$ $MgCO_3 \rightarrow MgO + CO_2$

5 Group 2 Elements

5.1 Trend and Properties of Group 2 Elements

LO (b) describe and explain qualitatively the trends and variations in atomic radius, ionic radius, first ionisation energy and electronegativity down a Group in terms of increasing number of electronic shells and nuclear charge

(f) describe and deduce from $E^\ominus$ values the relative reactivity of elements of Group 2 as reducing agents

<table>
<tr><th>Element</th><th>Atomic No.</th><th>Electronic Configuration</th><th>Atomic and ionic radius</th><th>First I.E.</th><th>Electronegativity</th><th>Melting point</th></tr>
<tr><td>Be</td><td>4</td><td>[He] $2s^2$</td><td colspan="3">Down the group,
⚛ Both nuclear charge and shielding effect increase
⚛ However, valence electrons are located in a shell with a larger principal quantum number, <i>n</i>, and are further away from the nucleus</td><td rowspan="5">Group 2 metals have giant metallic lattice structure with strong metallic bonds between the metal cations and sea of delocalised electrons.

Down the group,
⚛ Size of cations increases
⚛ Charge density decreases
⚛ Weaker electrostatic force of attraction between the cations and sea of delocalised electrons
⚛ Less energy is required to overcome the weaker metallic bonds
⚛ Melting point decreases</td></tr>
<tr><td>Mg</td><td>12</td><td>[Ne] $3s^2$</td><td rowspan="4">⚛ Weaker electrostatic forces of attraction between the nucleus and the valence electrons
⚛ Atomic and ionic radius increases</td><td rowspan="4">⚛ Weaker electrostatic forces of attraction between the nucleus and the valence electrons
⚛ Less energy needed to remove the valence electron
⚛ First I.E. decreases</td><td rowspan="4">⚛ Weaker electrostatic forces of attraction between the nucleus and the electron pair in a covalent bond
⚛ Electronegativity decreases</td></tr>
<tr><td>Ca</td><td>20</td><td>[Ar] $4s^2$</td></tr>
<tr><td>Sr</td><td>38</td><td>[Kr] $5s^2$</td></tr>
<tr><td>Ba</td><td>56</td><td>[Xe] $6s^2$</td></tr>
</table>

Group 2 element	Atomic No.	Electronic Configuration	Reducing power/ Chemical reactivity
Be	4	[He] $2s^2$	Down the group, ⚛ **Atomic radius increases** ⚛ Weaker **electrostatic force of attraction** between the nucleus and valence electrons ⚛ Metal atoms **lose their valence electrons** to form M^{2+} cations **more easily** ⚛ **Greater tendency to be oxidised and reducing power/ chemical reactivity increases down the group** *Evidence*: The **standard electrode potential, $E^\ominus$ value** becomes **more negative** down the group. $\underline{E^\ominus}$ $Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$ –2.37 V $Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$ –2.87 V $Sr^{2+}(aq) + 2e^- \rightleftharpoons Sr(s)$ –2.89 V $Ba^{2+}(aq) + 2e^- \rightleftharpoons Ba(s)$ –2.91 V More negative $E^\ominus$ value ⇒ greater tendency to be oxidised ⇒ stronger reducing power $M^{2+}(aq) + 2e^- \rightleftharpoons M(s)$ ⚛ Tendency for $M^{2+}(aq)$ to be ***reduced*** to M **decreases** ⚛ Tendency for Group 2 metals to be ***oxidised*** to M^{2+} **increases** ⚛ **Reducing power** of Group 2 metals **increases** ⚛ **Reactivity** of Group 2 elements **increases** (*e.g.* Ba is more reactive than Mg) *Note: Details about what the $E^\ominus$ value represents will be covered in the topic of Electrochemistry in JC2.*
Mg	12	[Ne] $3s^2$	
Ca	20	[Ar] $4s^2$	
Sr	38	[Kr] $5s^2$	
Ba	56	[Xe] $6s^2$	

5.2 Thermal Stability of Group 2 Oxysalts

LO (g) describe and explain the trend in thermal stability of Group 2 carbonates in terms of the charge density of the cation and the polarisability of the large anion

An oxyanion, or oxoanion, is an oxygen-containing polyatomic anion with the generic formula $A_xO_y^{z-}$, where A is another element. Oxyanions such as CO_3^{2-}, NO_3^-, SO_4^{2-}, IO_3^- and OH^- are generally **thermally unstable**, **decomposing on heating** to form **oxides**. The thermal stability of the Group 2 oxysalts is dependent on the following factors:

(i) size and charge of cation
(ii) size of oxyanion

Group 2 carbonates, nitrates and hydroxides decomposes in the following manner:

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

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

Hydroxides: $M(OH)_2\,(s) \rightarrow MO\,(s) + H_2O\,(g)$

Concept of Thermal Decomposition of Group 2 Carbonates

The **thermal stability** of Group 2 carbonates **increases** down the group. In other words, the **thermal decomposition temperature** of Group 2 carbonates **increases** down the group.

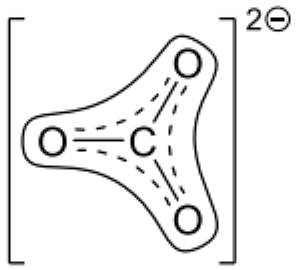
Down the group,

- **size of cations increases** (while charge remains unchanged)
- **charge density** (and polarising power) **of the cations decreases**
- **electron cloud of CO_3^{2-}** is **polarised (or distorted) to a lesser extent**
- **C–O bond** is weakened to a **lesser extent**.

Hence, thermal stability of group 2 carbonates **increases down the group**.

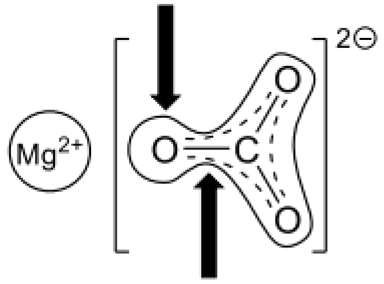
As seen from the figure on the right, group 2 carbonates with **cations of higher charge density** (e.g. Mg^{2+}) tend to undergo thermal decomposition **more easily** due to the **greater polarisation (or distortion) of the large CO_3^{2-} electron cloud,** which **weakens the C–O bond to a greater extent.**

Hence, **$MgCO_3$ is less thermally stable than $BaCO_3$.**



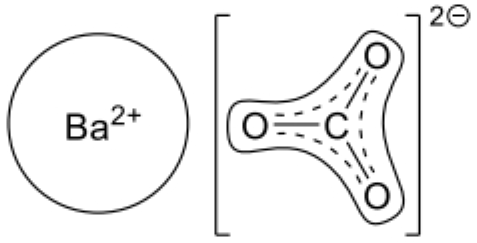
The electron cloud of CO_3^{2-} ion before polarisation.

oxygen atom on the way to becoming the O^{2-} anion



delocalised electrons being pulled towards the positive charge, weakening C–O bond

As the electron cloud of CO_3^{2-} ion is large, it is greatly polarised by the Mg^{2+} ion (of higher charge density).



The electron cloud of CO_3^{2-} ion is less distorted by the Ba^{2+} ion, as it has a lower charge density than Mg^{2+}.

Checkpoints for Section 5

- To describe and explain qualitatively the trends and variations in atomic radius, ionic radius, first ionisation energy and electronegativity down Group 2 in terms of increasing number of electronic shells and nuclear charge
- To describe and deduce from $E^\ominus$ values the relative reactivity of elements of Group 2 as reducing agents
- To describe and explain the trend in thermal stability of Group 2 carbonates in terms of the charge density of the cation and the ease of distortion (or polarisability) of the large anion, including writing of equations

Example 5A

1 Which factors help to explain the increase in thermal stability of the carbonates of Group 2 metals from magnesium to barium?

	charge density of cation	*relative polarisability of carbonate and oxide ions*
A	decreases	CO_3^{2-} less than O^{2-}
B	decreases	O^{2-} less than CO_3^{2-}
C	increases	CO_3^{2-} less than O^{2-}
D	increases	O^{2-} less than CO_3^{2-}

N03/I/16

2 *Use of the Data Booklet is relevant to this question.*
Lead(II) carbonate and zinc carbonate decompose on heating in the same way as calcium carbonate.
What will be the predicted order of decomposition temperatures of these three carbonates?

	highest				*lowest*
A	$CaCO_3$	>	$PbCO_3$	>	$ZnCO_3$
B	$CaCO_3$	>	$ZnCO_3$	>	$PbCO_3$
C	$PbCO_3$	>	$CaCO_3$	>	$ZnCO_3$
D	$ZnCO_3$	>	$CaCO_3$	>	$PbCO_3$

N02/I/15

6 Group 17 Elements

6.1 Trend and Properties of Group 17 Elements

Element	Structure and bonding	Boiling point/ Volatility $\left(\text{volatility} \propto \frac{1}{\text{boiling point}}\right)$	Colour			
			at r.t.p.	In gaseous state	In aqueous state	In organic solvent (*e.g.* hexane, CCl_4)
Cl_2	• **Simple molecular structure** • **Weak instantaneous dipole–induced dipole** (id-id) interactions between the **non-polar** X_2 molecules	Down the group, • **Number of electrons** increases • **Polarisability of electron cloud** increases • **More energy** required to overcome the **stronger intermolecular id–id interactions** • **Higher boiling point/lower volatility**	Greenish yellow **gas**	Greenish yellow	Pale yellow	Pale yellow
Br_2			Reddish brown **liquid**	Reddish brown	Orange	Orange-red
I_2			Black **solid**	Purple	Brown	Purple

LO (f) describe and deduce from $E^\ominus$ values the relative reactivity of elements of Group 17 as oxidising agents

Element	**Oxidising power and reactivity**	**Displacement reaction**
	Down the group, Size of atoms increases/ Atomic radius increases **Decrease** in tendency for X_2 to **accept electrons**. **Decrease** in tendency for X_2 to be **reduced** to X^-. **Oxidising power of halogens decreases** *Evidence*: The **standard electrode potential, $E^\ominus$ value** becomes **less positive** down the group. $E^\ominus$ $Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$ +1.36 V $Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$ +1.07 V $I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$ +0.54 V Less positive $E^\ominus$ value $\Rightarrow$ smaller tendency to be reduced $\Rightarrow$ weaker oxidising power *Note: Details about what the $E^\ominus$ value represents will be covered in the topic of Electrochemistry in JC2.*	Oxidising power of $Cl_2 > Br_2 > I_2$ The **more reactive halogens, X_2** (higher up in the group) can **displace the less reactive halides, X^-** (lower down the group) from their **aqueous solutions**.
Cl_2		Cl_2 can displace Br^- and I^-: $Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$ $Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$
Br_2		Br_2 can displace I^-: $Br_2 + 2I^- \rightarrow I_2 + 2Br^-$
I_2		I_2 is the **weakest** oxidising agent of all the halogens. It cannot displace Cl^- and Br^-.

Example 6A

When Cl_2 and Br_2 reacts with $Na_2S_2O_3$, Na_2SO_4 is formed. However, when I_2 reacts with $Na_2S_2O_3$, $Na_2S_4O_6$ is formed instead. Explain why this is so.

Cl_2 and Br_2 are **stronger** oxidising agents and hence can oxidise S in $S_2O_3^{2-}$ to a **greater** extent – oxidation number of S increases from **+2** in $S_2O_3^{2-}$ to **+6** in SO_4^{2-}.
I_2 is a **weaker** oxidising agent hence it oxidises S in $S_2O_3^{2-}$ to a **smaller** extent -- Oxidation number of S increases from **+2** in $S_2O_3^{2-}$ to **+2.5** in $S_4O_6^{2-}$.

6.2 Trend and Properties of Hydrogen Halides (HX)

LO (g) describe and explain the trend in thermal stability of Group 17 hydrides in terms of bond energies

<table>
<tr><th>Hydrogen halide</th><th>Structure and bonding</th><th>Boiling point</th><th colspan="2">Thermal stability</th></tr>
<tr><td>HCl</td><td rowspan="3">- Simple molecular structure
- Weak instantaneous dipole–induced dipole (id-id) and permanent dipole–permanent dipole (pd–pd) interactions between the polar HX molecules</td><td rowspan="3">Down the group,
- Number of electrons increases
- Polarisability of electron cloud increases
- More energy required to overcome the stronger id–id interactions
- Higher boiling point</td><td rowspan="3">When heated, $2HX(g) \rightarrow H_2(g) + X_2(g)$<table><tr><th>Bond</th><th>Bond Energy / kJ mol^{–1}</th></tr><tr><td>H–Cl</td><td>431</td></tr><tr><td>H–Br</td><td>366</td></tr><tr><td>H–I</td><td>299</td></tr></table>Down the group,
- Size of halogen atom increases
- Effectiveness of orbital overlap between H and X atoms decreases
- Bond strength of H–X decreases
- Bond energy of H–X decreases (<i>i.e.</i> H–X bond becomes weaker)
- Less energy required to break the H–X bond
- Thermal stability decreases</td><td>Does not decompose on heating.</td></tr>
<tr><td>HBr</td><td>Decomposes above 1000 °C.</td></tr>
<tr><td>HI</td><td>Decomposes readily on heating, <i>e.g.</i> plunging a red hot wire into the gas.</td></tr>
</table>

Checkpoints for Section 6

- To describe and deduce from $E^\ominus$ values the relative reactivity of elements of Group 17 as oxidising agents
- To describe and explain the trend in thermal stability of Group 17 hydrides in terms of bond energies

Example 6B

1 Group 17 elements and their hydrogen compounds exhibit trends in properties.

Which of these properties increases down the group C*l* → Br → I?

A the hydrogen-halogen bond dissociation energy

B the oxidising power of the elements

C the strength of the dispersion forces between halogen molecules

D the thermal stability of hydrogen halides

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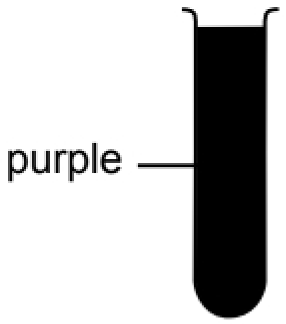
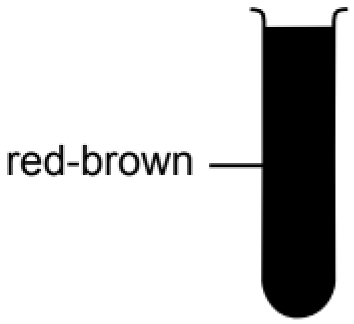
2 Aqueous chlorine is added to aqueous sodium bromide and the mixture is shaken with an equal volume of trichloroethane.
Which observation would be made?

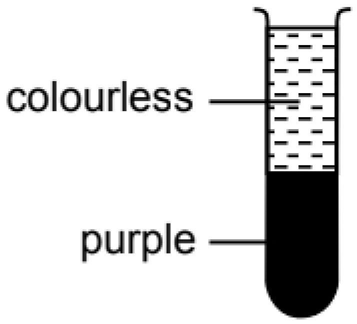
A



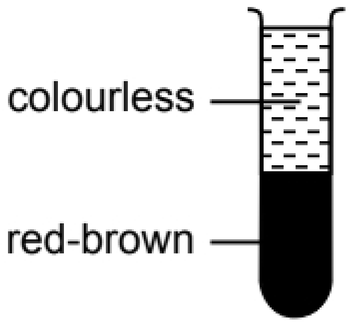
B



C



D



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Appendix 1 New Elements Added to the Periodic Table

From: Science Daily (https://www.sciencedaily.com/releases/2016/01/160130182456.htm)

Robert Grzywacz, along with collaborators at Oak Ridge National Laboratory, developed the software used in the equipment that detects the new elements and helps analyze data from the experiments. He is a co-author of recent papers, representing a collaboration between US and Russian scientists, that present new data on elements 113, 115, 117, and 118.

The International Union for Pure and Applied Chemistry recently announced formal verification of the four new chemical elements and recognized ORNL for the discovery of two: 115, temporarily named ununpentium (Uup, element 115), and 117, temporarily named ununseptium (Uus, element 117).

The four new elements complete the seventh row of the periodic table.

"The super-heavy element research is one of the most interesting efforts in nuclear physics," said Grzywacz, a professor in the UT Department of Physics and Astronomy. "It concerns the core question of how protons and neutrons form bound systems -- the nuclei. For us experimentalists, it is a formidable experimental challenge because the synthesis of super-heavy elements is incredibly difficult. It takes a concerted effort of many people to make it work. It is a fantastic but very also very time- and labor- intensive research program."

Grzywacz is director of the UT-ORNL Joint Institute for Nuclear Physics and Applications (JINPA). He was formerly an ORNL Wigner Fellow.

rzywacz and the ORNL team developed a data acquisition technology that uses a new type of digital signal processing to measure very fast nuclear decays down to a microsecond -- one millionth of a second. Because the system was relatively untested, he and former UT postdoctoral researchers David Miller and Nathan Brewer tested and debugged the system, which resulted in a reliable process. The data acquisition system initially applied in the ORNL-based experiments was already used in studies searching for and detecting super-heavy nuclei in laboratories in Dubna, Russia, and Darmstadt, Germany.

The ORNL team worked with scientists from Russia and the United States on the discovery of two of the new elements. Russian scientists used actinide targets and intense 48Ca beams to manufacture new super-heavy elements. The ORNL group supplied the target material to produce the new elements, which can be synthesized only at an ORNL reactor facility. The UT-ORNL collaboration recently provided a new, more sensitive detection system used in experiments in Russia.

The ORNL team will have the honor of helping name elements 115 and 117. New elements can be named after a mythological concept, a mineral, a place or country, a property, or a scientist.

At the moment, there is no direct application for the two new elements because the quantities that can be produced are very small and unstable against decay, Grzywacz said. The properties of these elements may help scientists predict more new elements that can be synthesized, with some of them predicted to be more stable.

Grzywacz noted that one of the beauties of research is that every theoretical claim has to be verified experimentally.

"Experiments produce surprises for theories; theories make unexpected predictions for experiments," he said. "This will never stop. There will be always an exciting problem to solve -- and by the way, sometimes, and often unexpectedly, the solution is actually useful for everyone."

Further readings:

Names and symbols of four newly discovered elements announced

Is It Time to Upend the Periodic Table?