‘A’ Levels H2 Chemistry - Periodic Table

This article is written based on the 9729 Chemistry GCE Advanced Level H2 Syllabus.
https://www.seab.gov.sg/docs/default-source/national-examinations/syllabus/alevel/2024syllabus/9729_y24_sy.pdf

The periodic table is one of the most powerful tools in chemistry, organising the elements in a way that reveals patterns in their properties and behaviors. For A-Level H2 Chemistry students, a deep understanding of the periodic table is essential, as it underpins much of the subject’s advanced concepts, from atomic structure to chemical bonding and reactivity trends.

In this article, let’s delve deeper into the trends and variations of atomic, physical, and chemical properties.

Trends and Variations in Electronic Configuration

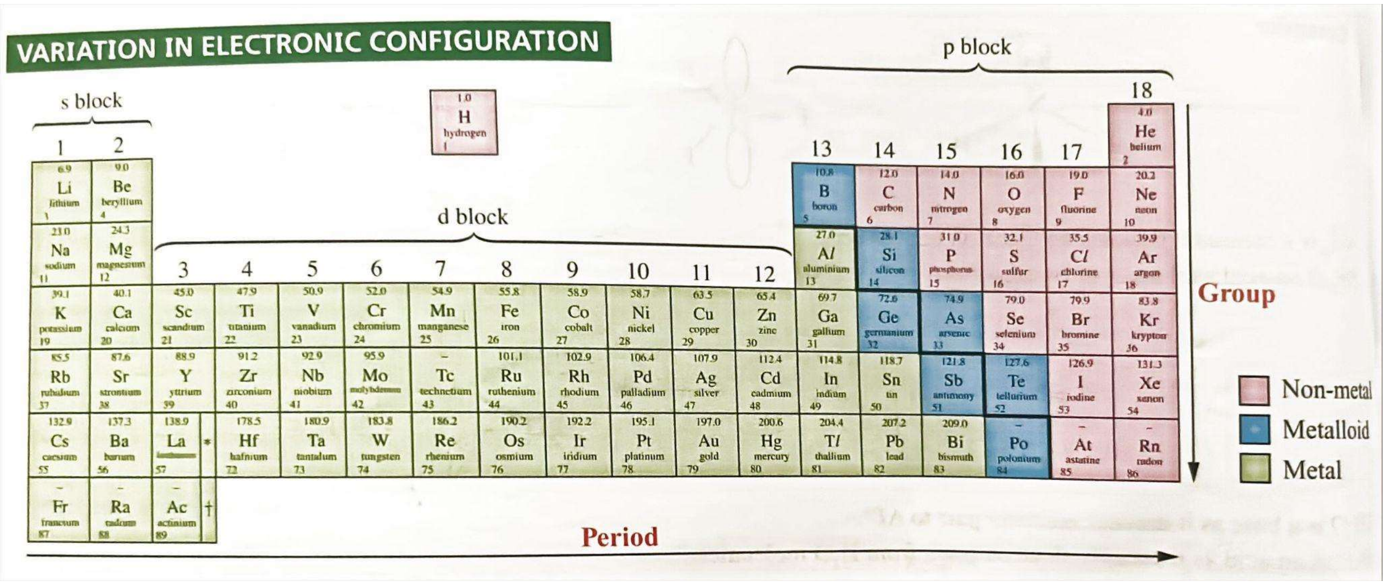
The modern periodic table is arranged in order of increasing atomic number (the number of protons in an atom's nucleus). The elements are organized into rows called **periods** and columns called **groups**.

Across a period, elements have

- Same number of outermost quantum shell
- Similar core electrons (e.g., Na: $\mathbf{1s^2 2s^2 2p^6} 3s^1$, A*l*: $\mathbf{1s^2 2s^2 2p^6} 3s^2 3p^1$)
- Periodic changes in physical and chemical properties

Down a group, elements have

- Same number of valence electrons (e.g., Li: $1s^2 \mathbf{2s^1}$, Na: $1s^2 2s^2 2p^6 \mathbf{3s^1}$)
- Similar physical and chemical properties



VARIATION IN ELECTRONIC CONFIGURATION

1.0
H
hydrogen
1

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
																	4.0 He helium 2
6.9 Li lithium 3	9.0 Be beryllium 4											10.8 B boron 5	12.0 C carbon 6	14.0 N nitrogen 7	16.0 O oxygen 8	19.0 F fluorine 9	20.2 Ne neon 10
23.0 Na sodium 11	24.3 Mg magnesium 12											27.0 A*l* aluminium 13	28.1 Si silicon 14	31.0 P phosphorus 15	32.1 S sulfur 16	35.5 C*l* chlorine 17	39.9 Ar argon 18
39.1 K potassium 19	40.1 Ca calcium 20	45.0 Sc scandium 21	47.9 Ti titanium 22	50.9 V vanadium 23	52.0 Cr chromium 24	54.9 Mn manganese 25	55.8 Fe iron 26	58.9 Co cobalt 27	58.7 Ni nickel 28	63.5 Cu copper 29	65.4 Zn zinc 30	69.7 Ga gallium 31	72.6 Ge germanium 32	74.9 As arsenic 33	79.0 Se selenium 34	79.9 Br bromine 35	83.8 Kr krypton 36
85.5 Rb rubidium 37	87.6 Sr strontium 38	88.9 Y yttrium 39	91.2 Zr zirconium 40	92.9 Nb niobium 41	95.9 Mo molybdenum 42	– Tc technetium 43	101.1 Ru ruthenium 44	102.9 Rh rhodium 45	106.4 Pd palladium 46	107.9 Ag silver 47	112.4 Cd cadmium 48	114.8 In indium 49	118.7 Sn tin 50	121.8 Sb antimony 51	127.6 Te tellurium 52	126.9 I iodine 53	131.3 Xe xenon 54
132.9 Cs caesium 55	137.3 Ba barium 56	138.9 La [illegible] 57 *	178.5 Hf hafnium 72	180.9 Ta tantalum 73	183.8 W tungsten 74	186.2 Re rhenium 75	190.2 Os osmium 76	192.2 Ir iridium 77	195.1 Pt platinum 78	197.0 Au gold 79	200.6 Hg mercury 80	204.4 T*l* thallium 81	207.2 Pb lead 82	209.0 Bi bismuth 83	– Po polonium 84	– At astatine 85	– Rn radon 86
– Fr francium 87	– Ra radium 88	– Ac actinium 89 †															

Several key periodic trends emerge from the arrangement of elements in the periodic table, and these trends are critical for understanding the chemical behavior of elements.

Trends and Variations in Atomic and Physical Properties

There are a few factors that commonly affect the atomic and physical properties of elements, such as atomic and ionic radius, electronegativity, and ionisation energy.

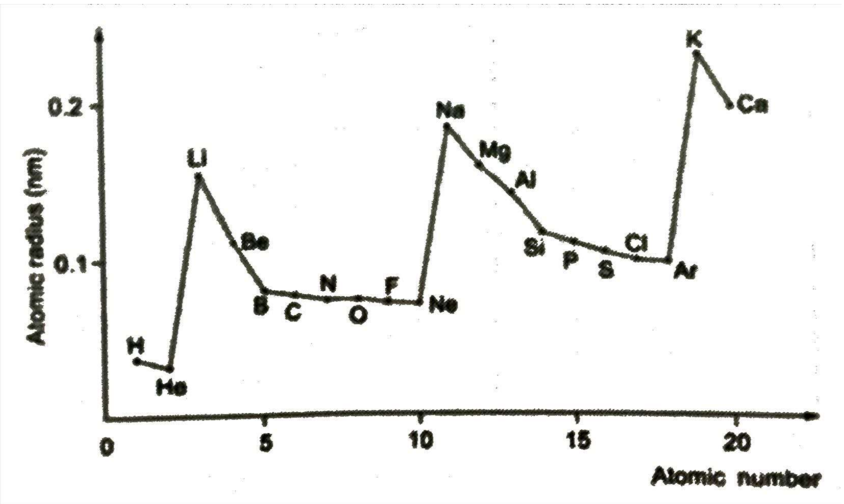
1) **Number of filled quantum shells**
2) **Nuclear charge, Z**
   - Positive charge of nucleus due to positively charged protons, which provide attraction on negatively charged electrons
3) **Shielding effect by inner shells of electrons**
   - Valence electrons shielded from attraction of nucleus by inner shells of electrons, thus full attractive force is not experienced

The net attractive force of the nucleus on the valence electrons, also termed as **effective nuclear charge, Z_{eff},** depends on both the nuclear charge and the screening effect.

$$Z_{eff} = Z - \text{screening effect}$$

Let's see how we can apply these factors to explain atomic and physical properties of elements across the periodic table.

Atomic Radius

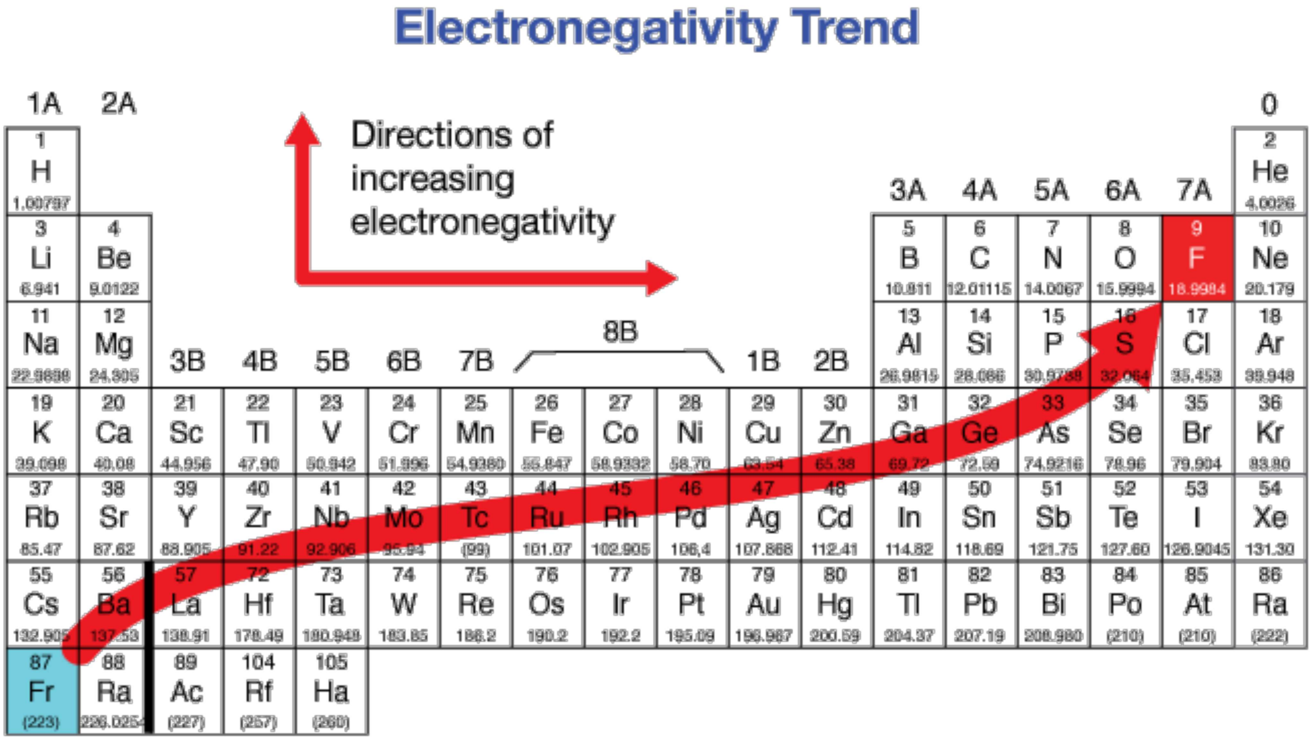


Down the group	Across the period
↑ number of protons, ↑ nuclear charge (Z)	↑ number of protons, ↑ Z
↑ number of inner shell of electrons, ↑ shielding effect	Number of inner shell of electrons is constant, shielding effect relatively constant
↑ screening effect outweighs ↑ Z as valence electrons are further away from the nucleus	↑ effective Z

↓ effective Z	
↓ strength of electrostatic attraction between nucleus and valence electrons	↑ strength of electrostatic attraction between nucleus and valence electrons
↑ atomic radius	**↓ atomic radius**

Electronegativity

Electronegativity is the relative tendency of an atom to **attract electrons** in a covalent bond.

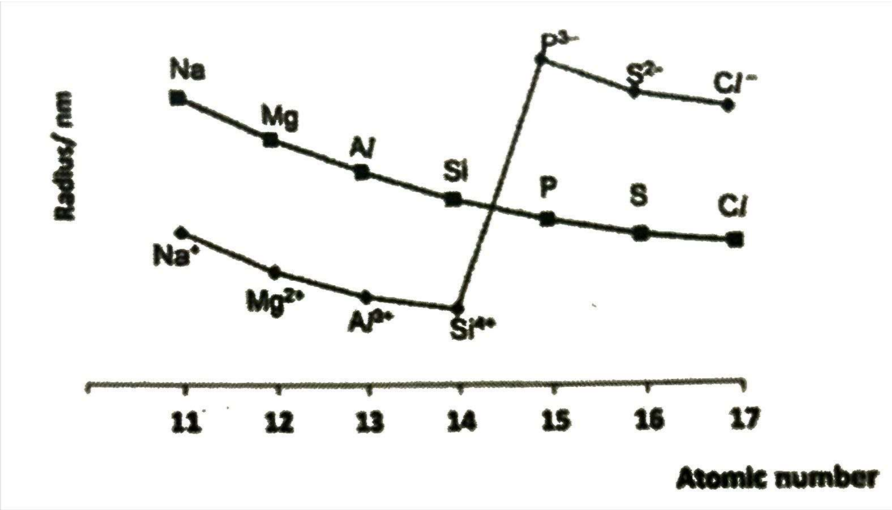


Source: https://chemistrytalk.org/electronegativity-chart-trends/

Down the group	**Across the period**
↑ number of protons, ↑ Z	↑ number of protons, ↑ Z
↑ number of inner shell of electrons, ↑ shielding effect	Number of inner shell of electrons is constant, shielding effect relatively constant
↑ screening effect outweighs ↑ Z as valence	↑ effective Z

electrons are further away from the nucleus	
↓ effective Z	
↓ strength of electrostatic attraction between nucleus and valence electrons	↑ strength of electrostatic attraction between nucleus and valence electrons
↑ atomic radius	**↓ atomic radius**
↓ tendency to attract a bonding pair of electrons	↑ tendency to attract a bonding pair of electrons
↓ electronegativity	**↑ electronegativity**

Ionic Radius



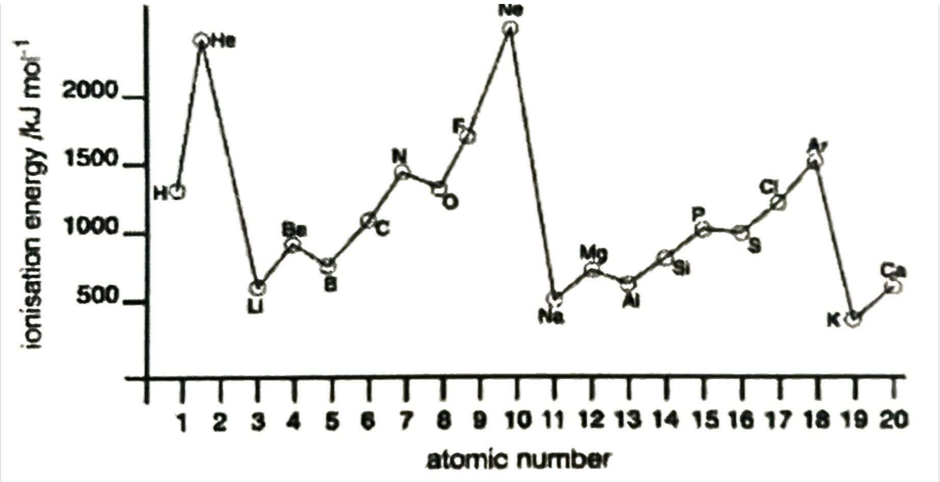
Cation	**Anion**
Atoms lose some/all valence electrons to form cations	Electrons added to outermost quantum shell of atoms to form anions
Cations have 1 less quantum shell of electrons than neutral atom	Increase electron-electon repulsion in outermost quantum shell of electrons
Ionic radii of cations SMALLER than atomic radii of corresponding atoms	**Ionic radii of anions LARGER** than atomic radii of corresponding atoms

Isoelectronic Series (e.g., Na^+ to Si^{4+}, P^{3-} to Cl^-)
↑ number of protons, ↑ Z
Number of inner shell of electrons is constant, shielding effect relatively constant
↑ effective Z
↑ strength of electrostatic attraction between nucleus and valence electrons
↓ ionic radius

First Ionisation Energy, IE

First Ionization Energy is defined as the amount of energy required to **remove one mole of electrons from one mole of gaseous atoms** to form one mole of gaseous singly charged positive ions.

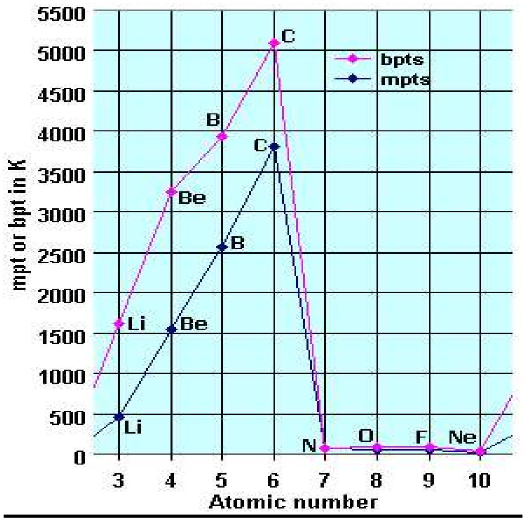
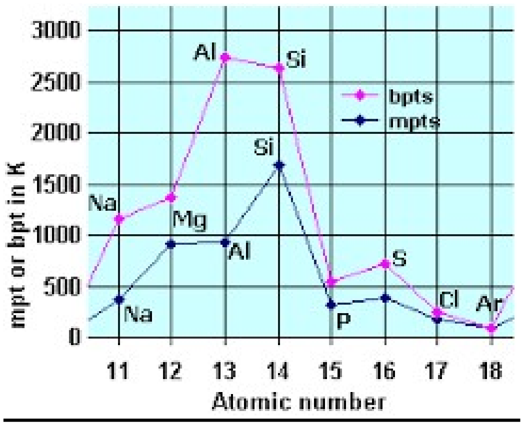
$$X(g) \rightarrow X^{+}(g) + e^{-}$$



Down the group	**Across the period**
↑ number of protons, ↑ Z	↑ number of protons, ↑ Z
↑ number of inner shell of electrons, ↑ shielding effect	Number of inner shell of electrons is constant, shielding effect relatively constant
↑ screening effect outweighs ↑ Z as valence electrons are further away from the nucleus	↑ effective Z
↓ effective Z	
↑ atomic radius	**↓ atomic radius**
↓ strength of electrostatic attraction between nucleus and valence electrons	↑ strength of electrostatic attraction between nucleus and valence electrons
↓ energy required to remove valence electrons	↑ energy required to remove valence electrons
↓ 1st IE	**↑ 1st IE**
	Take note of the following abnormalities: From Group 2 to 13 (e.g., Be to B - $1s^22s^2$ to $1s^22s^22p^1$)

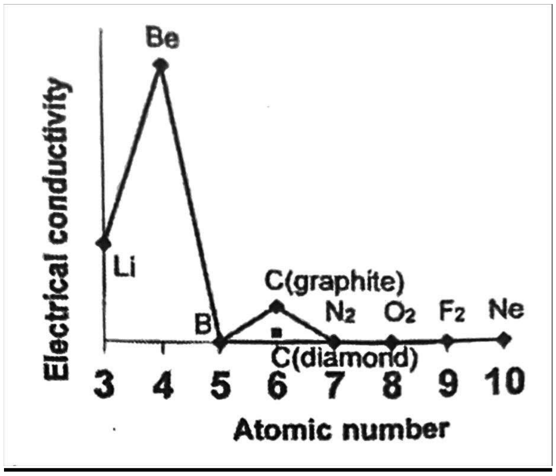
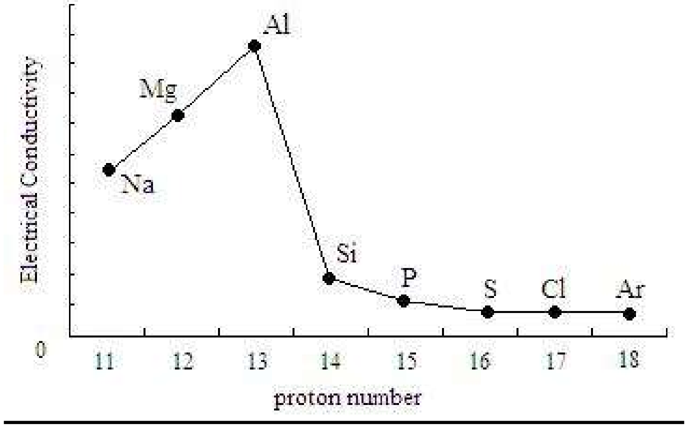
	Less energy required to remove a 2p electron than a 2s electron as the 2p electron is **further from nucleus** From Group 15 to 16 (e.g., N to O - $1s^22s^22p^3$ to $1s^22s^22p^4$) Less energy required to remove an electron from paired electron than unpaired electron due to **electrostatic repulsion between paired electrons**

Melting and Boiling Point

Across Period 2	**Across Period 3**
	
1) ↑ MP/BP from Li to Be - Li, Be have giant metallic structures - Large amount of energy required to overcome strong electrostatic attraction between positively charged metal ions and delocalised valence electron - ↑ number of delocalised valence electron - ↑ strength of metallic bonds - ↑ MP/BP	**1) ↑ MP/BP from Na to A*l*** - Na, Mg, A*l* have giant metallic structures - Large amount of energy required to overcome strong electrostatic attraction between positively charged metal ions and delocalised valence electron - ↑ number of delocalised valence electron - ↑ strength of metallic bonds - ↑ MP/BP

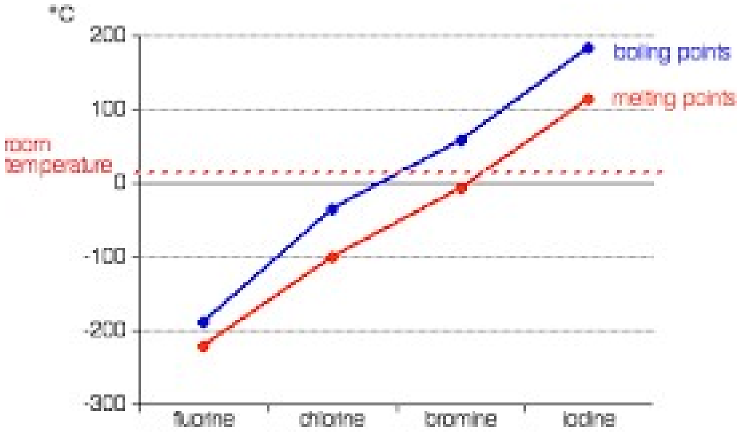
2) ↑ MP/BP from B to C - B, C have <u>giant covalent structures</u> - Large amount of energy required to overcome strong colavent bonds between atoms - ↓ size of bonding atom - ↑ effectiveness of orbital overlap - ↑ strength of covalent bond - ↑ MP/BP	**2) ↑ MP/BP to Si** - Si have <u>giant covalent structures</u> - Large amount of energy required to overcome strong colavent bonds between atoms - ↑ MP/BP
3) ↓ MP/BP sharply at N and remain low to Ne - N_2, O_2, F_2, Ne have <u>simple covalent structures</u> - Small amount of energy required to overcome weak instantaneous dipole-induced dipole interactions between discrete molecules - Low MP/BP	**3) ↓ MP/BP sharply at P and remain low to Ar** - P_4, S_8, Cl_2, Ar have <u>simple covalent structures</u> - Small amount of energy required to overcome weak instantaneous dipole-induced dipole interactions between discrete molecules - ↓ number of electrons - ↓ strength of id-id interactions - ↓ MP/BP

Electrical Conductivity

Across Period 2	Across Period 3
	
1) ↑ electrical conductivity from Li to Be - Li, Be have giant metallic structures - ↑ number of delocalised valence electron - ↑ electrical conductivity	**1) ↑ electrical conductivity from Na to A*l*** - Na, Mg, A*l* have giant metallic structures - ↑ number of delocalised valence electron - ↑ electrical conductivity
2) ↓ electrical conductivity sharply at 0 at B and remains at 0 to Ne (except C(graphite)) - B, C have giant covalent structures, N_2, O_2, F_2, Ne have simple covalent structures - Absence of mobile charged carriers to conduct electricity - B, C(diamond), N_2, O_2, F_2, Ne have no electrical conductivity	**2) ↓ electrical conductivity sharply at Si** - Si has giant covalent structures and is semi-conductor - Si has slight electrical conductivity
3) Abnomaly at C(graphite) - C(graphite) has giant covalent structure - Presence of delocalised electrons along layers to conduct electricity - C(graphite) has slight electrical conductivity	**3) Abnomaly at C(graphite)** - P_4, S_8, Cl_2, Ar have simple covalent structures - Absence of mobile charged carriers to conduct electricity - P_4, S_8, Cl_2, Ar have no electrical conductivity

Volatility

Volatility is the tendency to vapourise. It depends on the boiling point of substance.



Down Group 17
Halogens have simple covalent structure (exist as discrete molecules held together by weak instantaneous dipole-induced dipole interactions)
↑ number of electrons
↑ strength of instantaneous dipole-induced dipole interactions
↑ energy required to overcome interactions
↑ BP
↓ volatility

Trends and Variations in Chemical Properties

Apart from atomic and physical properties, there are also trends and variations in the chemical properties of the elements across the Periodic Table. Particularly, we are interested in the Period 3, Group 2, and Group 17 elements.

Chemical Properties of Chlorides of Period 3 Elements

Chloride	**$NaCl$**	**$MgCl_2$**	**$AlCl_3$**	**$SiCl_4$**	**PCl_5**
Chemical bonding	Ionic	Ionic (with covalent character)	Covalent (with ionic character)	Covalent	
Structure	Giant ionic		Simple covalent		
MP/BP	High		Low		
Oxidation Number	+1	+2	+3	+4	+5

Chloride	**Reaction with Water**		**pH**
$NaCl$	- **Hydration occurs** through ion-dipole interactions to form neutral solution - **Does not hydrolyse** in water (no reaction with water) due to low charge density of Na^+	Hydration: $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$	7
$MgCl_2$	- **Hydration occurs** through ion-dipole interactions - Undergoes **slight hydrolysis** in water to form slightly acidic solution due to higher charge density of Mg^{2+}	Hydration: $MgCl_2(s) + 6H_2O(l) \rightarrow [Mg(H_2O)_6]^{2+}(aq) + 2Cl^-(aq)$ Hydrolysis: $[Mg(H_2O)_6]^{2+}(aq) + H_2O(l) \leftrightarrow [Mg(H_2O)_5(OH)]^+(aq) + H_3O^+(aq)$	6.5
$AlCl_3$	- **Hydration occurs** through ion-dipole interactions - Undergoes **substantial hydrolysis** in water to form acidic solution due to higher charge density of Al^{3+} → able to polarise water molecules → weaken O-H bonds in water molecules → produce H_3O^+	Hydration: $AlCl_3(s) + 6H_2O(l) \rightarrow [Al(H_2O)_6]^{3+}(aq) + 3Cl^-(aq)$ Hydrolysis: $[Al(H_2O)_6]^{3+}(aq) + H_2O(l) \leftrightarrow [Al(H_2O)_5(OH)]^{2+}(aq) + H_3O^+(aq)$	3
$SiCl_4$	- **Hydration does not occur** - Undergoes **complete hydrolysis** in water to form very acidic solution	Hydrolysis: $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(aq)$	2

PCl_5	- **Hydation does not occur** - Undergoes **complete hydrolysis** in water to form very acidic solution	Hydrolysis: $PCl_5(l) + 4H_2O(l) \rightarrow H_3PO_4(s) + 5HCl(aq)$	2

Chemical Properties of Oxides of Period 3 Elements

Oxide	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_4O_{10}**	**SO_3**
Chemical bonding	Ionic		Ionic (with covalent character)	Covalent		
Structure	Giant ionic			Giant covalent	Simple covalent	
MP/BP	High				Low	
Type of oxides	Basic		Amphoteric	Acidic		

Oxide	**Reaction with Water**		**pH**
Na_2O	**Readily dissolves** in water to give alkaline solution	$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$	13
MgO	**Dissolves slightly** in water to give a weak alkaline solution	$MgO(s) + H_2O(l) \leftrightarrow Mg(OH)_2(aq)$	9
Al_2O_3	**Insoluble** in water due to high lattice energy	-	-
SiO_2	**Insoluble** in water due to high lattice energy	-	-
P_4O_{10}	**Readily dissolves** in water to give an acidic solution	$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$	2
SO_3	**Readily dissolves** in water to give an acidic solution	$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$	2

Oxide	**Reaction with Acid/Base**	
Na_2O	- **React with acid** to form salt and water - **Does not react with alkali**	$Na_2O(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l)$

MgO	- **React with acid** to form salt and water - **Does not react with alkali**	$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$
Al_2O_3	- **React with BOTH acid and alkali** to form salt and water	$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$ $Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$
SiO_2	- **React with concentrated alkali** to form salt and water	$SiO_2(s) + 2NaOH(aq) \rightarrow Na_2SiO_3(aq) + H_2O(l)$
P_4O_{10}	- **React with alkali** to form salt and water	$P_4O_{10}(s) + 12NaOH(aq) \rightarrow 4Na_3PO_4(aq) + 6H_2O(l)$
SO_3	- **React with alkali** to form salt and water	$SO_3(g) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

Chemical Properties of Group 2 Elements

Chemical Property	**Trend**	**Explanation**
Reducing Power	Increases down the group	- ↑ number of protons, ↑ Z - ↑ number of inner shell of electrons, ↑ shielding effect - ↑ screening effect outweighs ↑ Z as valence electrons are further away from the nucleus - ↓ effective Z - ↑ **atomic radius** - ↓ strength of electrostatic attraction between nucleus and valence electrons - ↓ energy required to remove valence electrons - ↓ **1st IE** - ↑ tendency of metal to be oxidised - $E^\ominus(M^{2+}/M)$ more negative - ↑ **reducing power**
Thermal Stability of Group 2 Compounds	Increases down the group	- ↑ **catonic radii** of M^{2+} - ↓ charge density (charge to size ratio) - ↓ ability to polarise electron cloud of large anion (e.g., CO_3^{2-}) - Bonds weakened to smaller extend - ↑ **thermal stability**

Chemical Properties of Group 17 Elements

Chemical Property	Trend	Explanation
Oxidising Power	Decreases down the group	- ↑ number of protons, ↑ Z - ↑ number of inner shell of electrons, ↑ shielding effect - ↑ screening effect outweighs ↑ Z as valence electrons are further away from the nucleus - ↓ effective Z - ↓ strength of electrostatic attraction between nucleus and valence electrons - **↑ atomic radius** - ↓ tendency to attract a bonding pair of electrons - **↓ electronegativity** - ↓ ease of gaining electrons to form halide ions during oxidation - $E^{\ominus}$ less positive - ↓ **oxidising power**
Thermal Stability of Hydrogen halides	Decreases down the group	- **↑ atomic radius** - ↑ diffusion of valence orbitals - ↓ effectiveness of orbital overlap between smaller H and larger X atoms - ↓ energy required to break H-X bond - **↓ thermal stability**

Final Comments

'Periodic Table' is a heavy topic in 'A' Levels H2 Chemistry, incorporating content from across different topics. However, it is definitely not difficult once you are able to understand the main group and period patterns associated with the periodic table. Do remember to practice with examples and regularly revisit periodic trends to reinforce your understanding. Best of luck, and may your enthusiasm for chemistry continue to inspire you in all your future pursuits!

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