

3A – The Periodic table

History of the Periodic table:

- Research on of the following who contributed to the development in ordering the elements.
- Put your findings in a PowerPoint presentation to present to the rest of the class:

Aristotle

Antoine - Laurent de Lavoisier

Jons Jakob Berzelius

Johan Wolfgang Dobereiner

Alexandre-Emile Beguyer de Chancoutois

John Newlands

Dmitri Mendeleev

Henry Moseley

Glenn Seaborg

The 4 elements

The list of elements

Symbols and Atomic weights

Law of Triads

Spiral cylinder / aligning similar elements

Law of Octaves

Atomic masses / Periodic law

Atomic number

Transuranic elements 94 - 102

The Periodic table: Mendeleev and beyond

- The Periodic Table is the Bible to Chemists. It brings order and a systematic way of looking at the elements.
- Prior to the periodic table, it was very difficult to find patterns in the elements.
- Although many people contributed to the development of the modern Periodic table, it falls mainly down to a handful:

Dobereiner – Law of Triads:-

- If you look at the properties and relative atomic masses of 3 elements in group 1:-

Element	RAM	Properties with water
Lithium	6.9	Fizzes gently – hydrogen given off
Sodium	23	Fizzes – hydrogen given off
Potassium	39.1	Fizzes violently – hydrogen given off

- The RAM of sodium is the mean of lithium and potassium.
- The rate of reaction of sodium is also in between the other 2.
- This also occurs for other 'triads' of elements.

Newlands – Law of Octaves:-

- Newlands was the first to notice that if you arrange the elements in order of RAM, every 8th element exhibited similar physical and chemical properties.
- This only worked for the first 16 elements and other elements were missing.

Na	Mg	Al	Si	P	S	Cl	Ar	K
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Groups:

- **A vertical column is called a group:**- Groups are numbered I - VII, and 0. Some groups have names as well.
- Groups often show an increase or decrease in similar properties.

Periods:

- **A horizontal row is called a period:**- Periods are numbered from the top downwards (1-7).
- Periods often show gradual changes in properties.
- These patterns are repeated across each Period.
- This is called **periodicity**

Periodicity:

Is a regular periodic variation of properties of elements with atomic number and the position in the Periodic Table

- Trends in properties can occur **across periods and down groups:**

Across a Period: Metal → Non – metal

Down Group 4: Non - Metal → Metal

Variation in electron structure:

- Chemical reactions are due to the outer shell electrons only.
- This means that those elements with the same number of electrons in the outer shell will react in similar ways.
- The vigour of the similarities will depend on shielding and the number of protons in the nucleus (Li - Cs).
- This is what causes Periodicity - the regular repeating patterns:

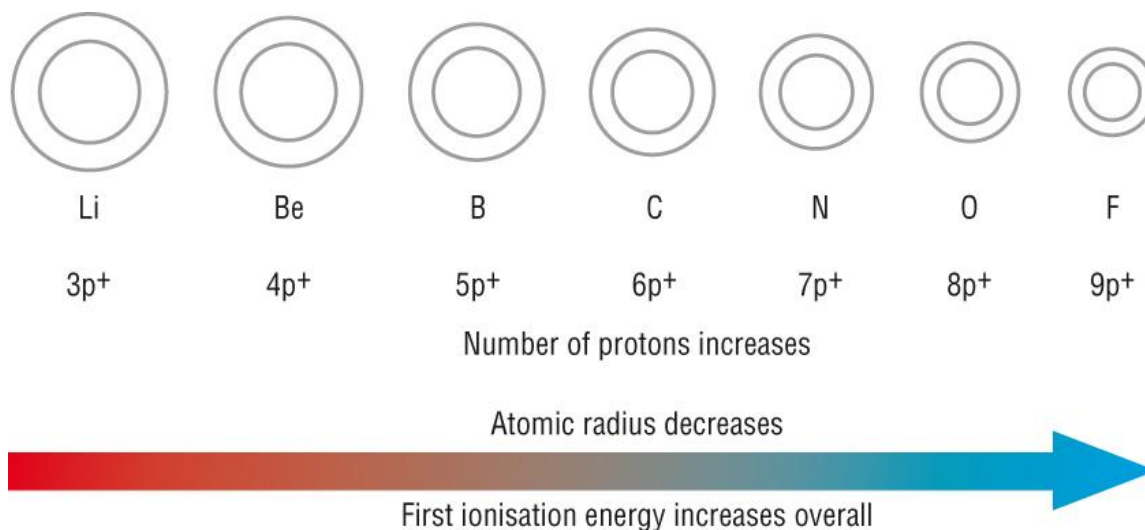
Li	Be	B	C	N	O	F	Ne
$[\text{He}]2s^1$	$[\text{He}]2s^2$	$[\text{He}]2s^22p^1$	$[\text{He}]2s^22p^2$	$[\text{He}]2s^22p^3$	$[\text{He}]2s^22p^4$	$[\text{He}]2s^22p^5$	$[\text{He}]2s^22p^6$
Na	Mg	Al	Si	P	S	Cl	Ar
$[\text{Ne}]3s^1$	$[\text{Ne}]3s^2$	$[\text{Ne}]3s^23p^1$	$[\text{Ne}]3s^23p^2$	$[\text{Ne}]3s^23p^3$	$[\text{Ne}]3s^23p^4$	$[\text{Ne}]3s^23p^5$	$[\text{Ne}]3s^23p^6$

Periodicity: Ionisation energies and atomic radii

A) Trends across a period:

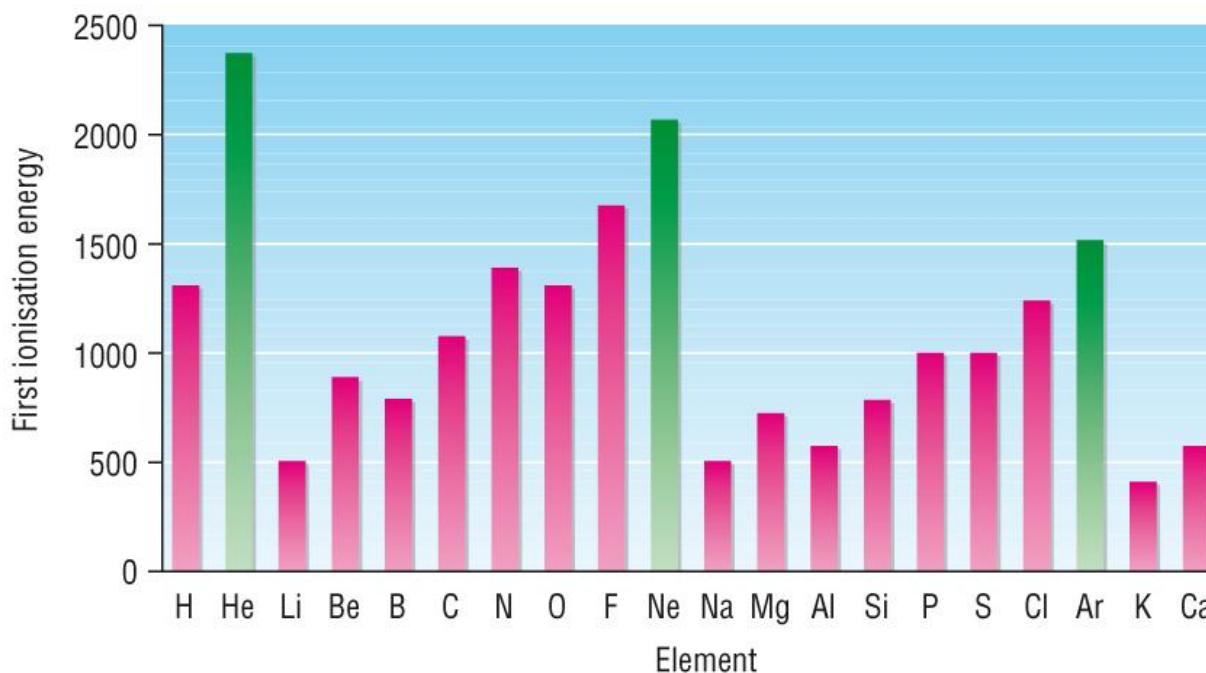
1) Periodic Patterns: Atomic radii

- As you go across a period, the **nuclear charge increases**.
- And the shielding between nucleus and outer electron shell remains the same.
- The increased nuclear charge pulls the electron shell closer to the nuclei.
- This means there is a slightly greater attraction between the nucleus and the outer shell electrons.
- **The increase in nuclear charge is the most important factor.**



2) Periodic Patterns: First ionisation energies

- The explanation for atomic radii explains the trend in first ionisation energies across a period
- Graph of 1st ionisation energies for the first 36 elements:



Recap – Factors affecting ionisation energies:

- 1) The distance of the electron from the nucleus
- 2) Size of the positive nuclear charge
- 3) The 'shielding' effect by full inner shells

Explanation:

Start with atomic radii:

- As you go across a period, the nuclear charge increases.
- As the distance between nucleus and outer shell remains the same
- And as shielding between nucleus and outer electron shell remains the same.
- The increased nuclear charge pulls the electron shell closer to the nuclei.
- This means there is a slightly greater attraction between the nucleus and the outer shell electrons.

Then include what effect that has on removing an electron:

- This means that the first ionisation energy increases across a Period.

B) Trends down a Group:

1) Down a Group: Atomic radii increases:

Number of shells increases

Shielding increases

Atomic radius increases

First ionisation energy decreases

- As you go down a group, the nuclear charge increases.
- There are more electron shells which increases **shielding** between nucleus and outer electron shell.
- As there are more electron shells, the **distance** between nucleus and outer shell increases.
- This means that attraction between the nucleus and outer electrons decreases.
- **Distance and shielding are the most important factors.**

2) Down a Group: First ionisation energies decreases:

Explanation:

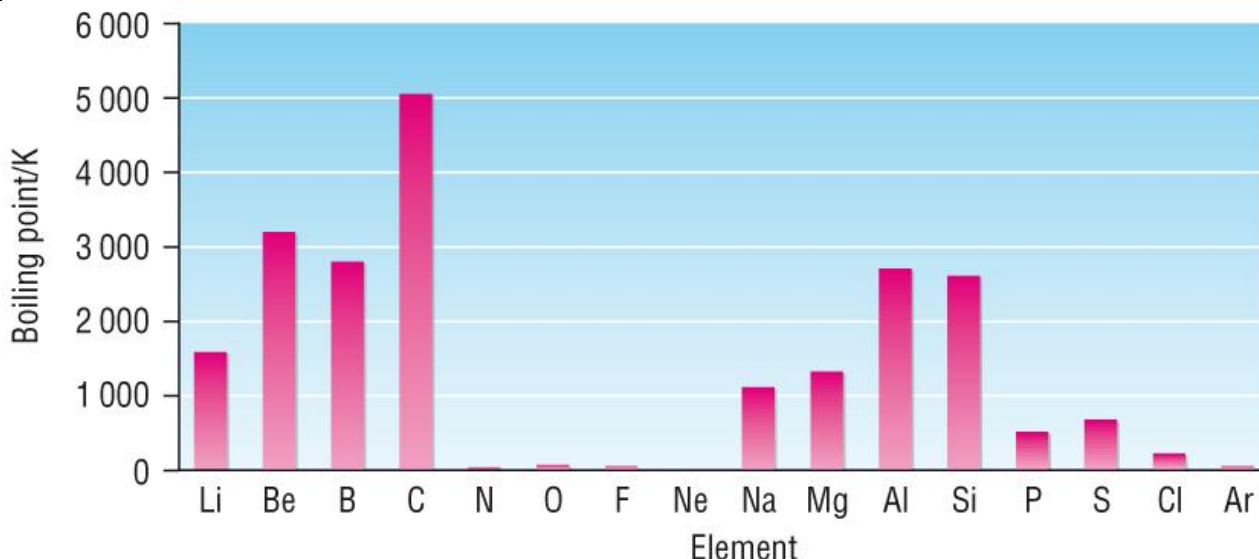
Start with atomic radii:

- As you go down a group, the nuclear charge increases.
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- This means that attraction between the nucleus and outer electrons decreases.

Then include what effect that has on removing an electron:

- This means that the first ionisation energy decreases down a Group.

Periodicity: Boiling points



General trends:

- Increase in Boiling point from Gp 1 - 4.
- Sharp drop from Gp 4 - 5
- Low Boiling points for Gp 5 - 0

Explanation:

- As you move from left to right there is a general pattern:

Metals → Non - metals

- And:

Solid → Gas

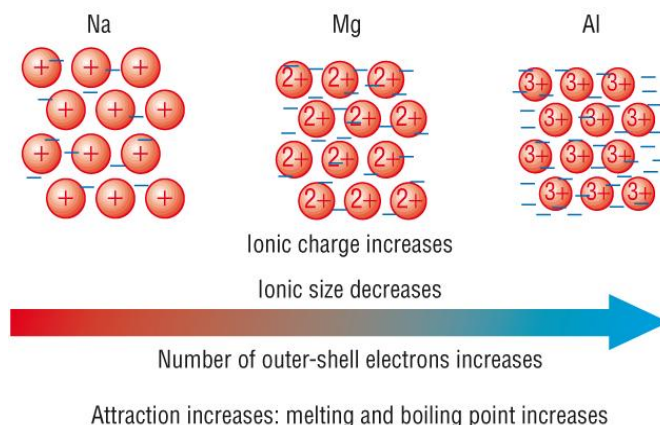
- This means that there are different **structures, forces of attraction** and **bonding**.
- All 3 of these must be considered when explaining boiling / melting points. As we move across a period.

Metallic → Giant Molecular → Molecular → Atomic

- If you understand the **structures, forces of attraction** and **bonding** in all 4 of the above, then it becomes easier:
- We shall take each one in turn:

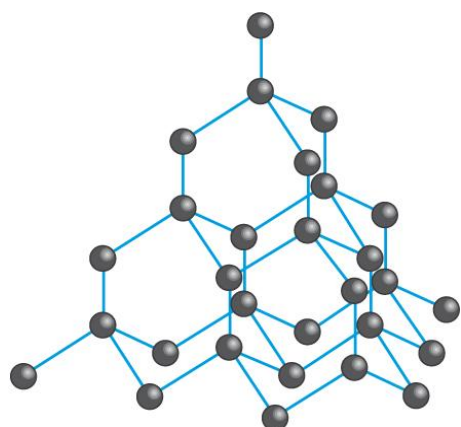
Metallic bonding - Gp 1 - 3

- The boiling points increase as you move along Group 1 - 3.
- This is due to an increase in outer electrons available to be part of the mobile sea of electrons.

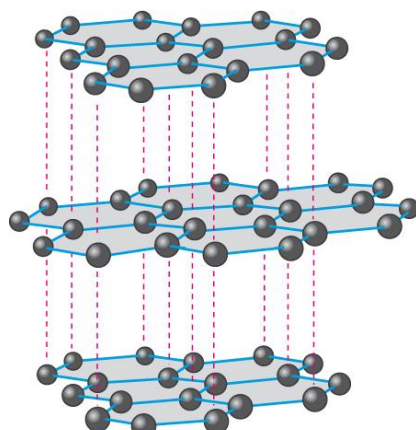


Giant covalent structures - Gp 4

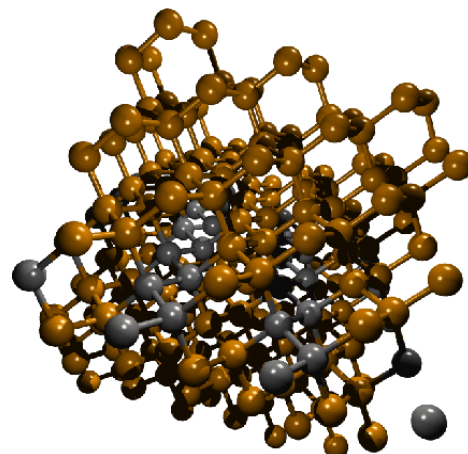
- These are carbon and silicon
- This is due to the strong covalent bonds making up a giant structure.
- For these to boil, the strong covalent bonds must be broken leaving single atoms.



Carbon - Diamond



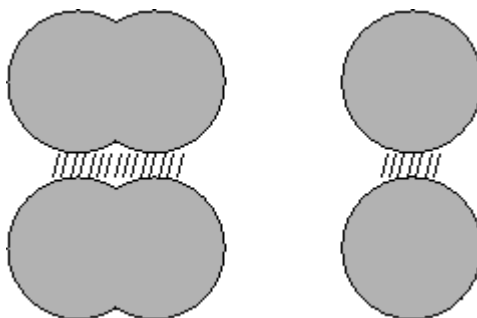
Carbon - Graphite



Silicon

Molecular / Atomic - Groups 5 - 0:

- These are elements that are diatomic molecules (N_2 , O_2) or single atoms (Ne, Ar).
- The intermolecular forces between these elements are Van der Waals and therefore weak.
- These Van der Waals are easily overcome separating the diatomic molecules or atoms from each other.



Summary:

Period 2	Li	Be	B	C	N_2	O_2	F_2	He
Period 3	Na	Mg	Al	Si	P_4	S_8	Cl_2	Ar
Structure	Giant metallic			Giant covalent	Simple molecular / atomic structures			
Forces	Strong forces between positive ions and negative delocalised electrons			Strong forces (bonds) between atoms	weak forces between molecules / atoms			
Bonding	Metallic bonding			Covalent bonding	Van Der Waals'			