Chem Factsheet

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Revision Summary: Trends in the Periodic Table 1

Before working through this Factsheet you should:
- Be familiar with the organisation, layout and information available from the Periodic Table;
- Have studied the AS and A2 work on periodic trends.

After working through this Factsheet you will have revised:
- The layout of the Periodic Table;
- Trends in physical properties and their explanation.

This Factsheet provides a revision summary of periodic table work. If you would like more explanations on any aspect, these can be found in Factsheets 12, 13, 14, 19 and 20. Factsheet 68 also looks at periodic trends in a different context.

Questions on trends in the Periodic Table are popular at AS and A2 level. Such questions often require two styles of preparation:

- Factual recall – statement of a trend and evidence of it, e.g. the recall of particular chemical equations.
- Application and understanding – answers require a description of **why** there is a particular trend.

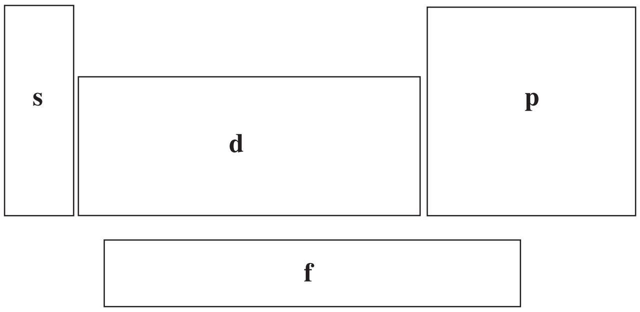
First, revision of the Periodic Table itself.

The elements are placed in order of increasing atomic number.

Period: Horizontal row of elements.
Group: Vertical column of elements.

s, p, d and f blocks

The Periodic Table is divided up into 4 blocks (s, p, d and f) which show which is the outer electron sub-shell for the elements in that block.



s-block elements: - the metals in group 1 and 2, so called because their outer shell contains s electrons.
e.g. Sodium (Na): Atomic no. 11. Elec. config. $1s^2 2s^2 2p^6 3s^1$

p-block elements: - the elements from groups 3 to 7, as they have outer electrons which are p-electrons.
e.g. Carbon (C): Atomic no. 6. Elec. config. $1s^2 2s^2 2p^2$

d-block elements: - the metals in the block between groups 2 and 3, known as the **transition metals**. d-block elements have incomplete d-sub shells.
e.g. Titanium (Ti): Atomic no. 22. Elec. config. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

f-block elements: - a block of elements within the transition metals, so called because electrons are being added into the f-subshell in these elements. e.g. Cerium (Ce) Atomic no. 58.
Elec. config. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4s^6 5s^2 4d^{10} 5p^6 6s^2 4f^2$

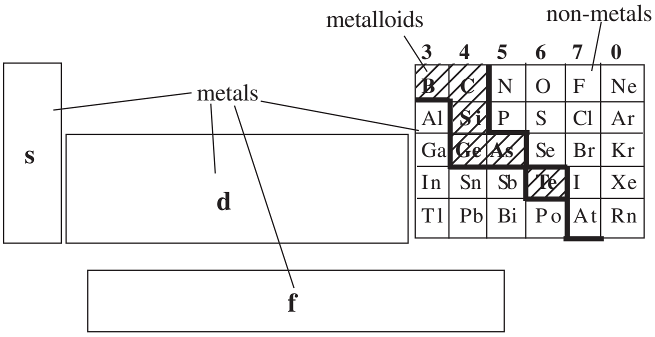
*Properties of an element are a periodic function of their atomic numbers. Properties repeat regularly, so that elements in the same **group** tend to display **similar chemical and physical properties**.*

You need to learn the following group names:

Group Number	Group Name
1	Alkali Metals
2	Alkaline Earth Metals
7	Halogens
0	Noble Gases

Metals, Metalloids and Non-metals

These can also be shown to be grouped on the Periodic Table:



To categorise the elements into one of these three types, the following criteria can be used:

Metals
- Good conductors of electricity
- Form basic oxides (a few form amphoteric oxides)

Metalloids
- Poor conductors of electricity (except graphite)
- Form amphoteric oxides

Non-metals
- Virtually non-conductors of electricity (insulators)
- Form acidic oxides

Note: You may find differences between textbooks on which elements are classified as "metalloids" and which as "non-metals". Do not worry about this - the key point is understanding the characteristics.

Trends in physical properties - Period 3

Group	1	2	3	4	5	6	7	0	**Notes**
Element	sodium	magnesium	aluminium	silicon	phosphorus	sulphur	chlorine	argon	
Symbol	Na	Mg	Al	Si	P	S	Cl	Ar	
Structure	← Giant Metallic →			Giant Covalent	← Molecular →			Atomic	*All metals have metallic structure - this leads to good conductance.* *Silicon has a (covalently bonded) diamond structure.* *Chlorine and the other halogens all exist as diatomic molecules; phosphorus and sulphur are simple molecular, but not diatomic. All noble gases are monatomic*
Character	← Metallic →			Metalloid/ Non-metal	← Non - Metallic →				*Metallic character generally decreases across a period; this is associated with the decrease in ionisation energy (so less likely to form positive ions) and increase in electronegativity/electron affinity (more likely to attract electrons)*
Melting Point	General increase				General decrease				*Melting point is determined by structure - for metals, generally the more valence electrons, the stronger the metallic bonding and the higher the melting point. Giant covalent structures always produce high melting points. For simple molecular structures, melting point is determined by the strength of van der Waals forces, which depend on molecular size*
Conductance	General increase			General decrease					*Associated with metal/metalloid/non-metal character - within the metals, the higher the number of valence electrons available, the higher the conductivity*
First Ionization Energy	General increase								*Note this just refers to forming the M^+ ion, not the commonest ion formed. It is always endothermic* *The effective nuclear charge increases across a period - since electrons are being added to the same shell, so there is comparatively less shielding. This makes it harder to remove an electron.* *Note the exceptions due to half-shell stability, full-sub-shell stability and p electrons being easier to remove than s-electrons - the 1st IE for aluminium is less than for magnesium, and for sulphur less than for phosphorus*
Atomic Radius	General decrease								*Again, this decreases due to increasing effective nuclear charge - electrons are held more tightly by the nucleus*
Ionic Radius	General decrease			Not applicable	General decrease			Not applicable	*This refers to the commonest ion formed - eg Na^+, S^{2-} etc* *For positive ions, the higher the charge, the smaller the ion, since the remaining electrons are pulled closer to the nucleus.* *For negative ions, the higher the charge, the larger the ions, since the nucleus is having to retain more additional electrons* *Silicon and argon do not form ions*
Ionic Radius compared to Atomic Radius	← Smaller →			Not applicable	← Larger →			Not applicable	*Positive ions are always smaller than the neutral atom, and negative ions larger, since the fewer the electrons, the closer they are pulled to the nucleus*
Electronegativity	General increase							Not applicable	*Measures the ability to attract electrons within a bond.* *Increases due to increasing effective nuclear charge.* *Bond type is determined by the difference in electronegativity, and it also helps explain the transition from metallic to non-metallic character*
First Electron Affinity	General increase							Not applicable	*Measures the ability to attract an electron - always exothermic* *Increases due to the increase in effective nuclear charge. Elements with high electron affinity are more likely to form negative ions. Note that this is not a* ***uniform*** *increase - atoms with full sub-shell or half sub-shell stability have a lower electron affinity than would be expected.*

Group trends

The table below shows **general trends** that apply throughout all groups of the periodic table; other trends are specific to particular groups.

General Trends

atomic radius	ionic radius	1st ionisation energy	Electronegativity	1st electron affinity
Increases ↓	Increases ↓	Decreases ↓	General Decrease ↓	General Decrease ↓

Notes

- *Atomic and ionic radius increase because of increasing number of shells of electrons*
- *First ionisation energy increases because the electron being removed is further from the nucleus and the increased shielding due to inner shells*
- *Electronegativity and electron affinity decrease due to the increased shielding.*
- *There are some anomalies in the decreasing trend for electronegativity -eg Group 4*
- *There are also anomalies in the decreasing trend for electron affinity - fluorine, for example, has a lower first electron affinity than chlorine, due to repulsion from the electrons already in place.*

Group 1 and 2 Trends

Members of groups 1 and 2 are all metals They are silver in colour and tarnish in air. Mant of their characteristic trends in their physical properties can be explained by the fact that they show weak metallic bonding.

melting point	boiling point	density
Decreases ↓	Decreases ↓	Increases ↓

Notes

The melting point decreases down the group because in the metallic bond, the number of delocalised electrons remains the same and the charge on each metal cation stays the same, but the ionic radius increases so the attraction between the delocalised electrons and the metal cations decreases.

The density increases because the atomic mass increases faster than the size of the atom.

Group 7 trends

Members of group 7 are all non-metals and have low melting and boiling points

melting point	boiling point
Increases ↓	Increases ↓

Notes

The melting and boiling points here are determined by the strength of the Van der Waals forces.

All the halogens form diatomic molecules, so as the group is descended, the molecules get larger. This produces larger Van der Waals forces, and hence increased melting and boiling points

Practice Questions

1. State and explain how the following properties vary for the elements of Period 3 (Na – Ar).
   (a) Atomic radius. (b) Melting point. (c) First ionisation energy.

2. (a) Explain why the ionic radius of bromine is larger than its atomic radius.
   (b) Explain why the ionic radius of Al^{3+} is smaller than that of Na^+

3. Electrical conductivity is measured in units called *siemens per metre*, $S\ m^{-1}$. Values for the elements in Period 3 are given in the table below.

Element	Electrical conductivity / $10^8\ S\ m^{-1}$
Na	0.218
Mg	0.224
Al	0.382
Si	10^{-10}
P	10^{-17}
S	10^{-23}
Cl	–
Ar	–

(a) Explain why the electrical conductivity of sodium, magnesium and aluminium is relatively high.

(b) Why does electrical conductivity increase from Na to Mg to Al?

(c) The electrical conductivity of boron is 10^{-12} (in units of $10^8\ S\ m^{-1}$) Use this value to state and explain the type of bonding present in boron

4. State and explain how the following properties vary going down a group of the Periodic Table
   (a) Atomic radius. (b) First ionisation energy.

5. In groups 1 and 2 of the Periodic Table, melting points **decrease** down the group, but in group 7, they **increase**. Explain this difference.

Answers

1. (a) Decreases across the period.
   Increasing nuclear charge, whilst electrons being added into the same quantum shell.
   (b) Increase Na – Si, then decrease P – Ar.
   Na – Si increasing metallic bond strength, Si very strong covalent bonding, P – Ar held by much weaker Van der Waals forces.
   (c) Increases across period.
   Increasing nuclear charge.

2. (a) Bromine forms Br^- ion, involving the addition of an electron.
   This causes electron repulsion to increase, increasing the size of the particle.
   (b) Al^{3+} and Na^+ have the same number of electrons, but Al^{3+} has a greater nuclear charge.
   Hence the electrons in Al^{3+} are more tightly held, making it a smaller particle

3. (a) They have metallic bonding, so electrons are delocalised and can move under an applied p.d. throughout the entire metal
   (b) Number of outer shell electrons increases from 1 to 2 to 3
   Hence there are more free / delocalised / mobile electrons
   (c) Covalent bonding. Electrical conductivity is low, so very limited movement of electrons

4. (a) Increases, due to increasing number of shells of electrons
   (b) Decreases, due to increased distance of electron from nuclues and increased shielding.

5. In groups 1 and 2, melting point is determined by strength of metallic bonding. This decreases, because although the charge on the cation and the number of delocalised electrons remain constant, the ionic radius increases, producing decreased attraction between the cations and electrons

   In group 7, melting point is determined by the strength of the Van der Waals forces; these increase, due to increase in molecular size.

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