Chem Factsheet

January 2001 Number 12

Periodicity - Trends in Period 3

To succeed in this topic you need to:-
- understand the work in Factsheet 1 on Atomic Structure
- understand the work in Factsheet 6 on structure of elements and compounds, including how bonding relates to physical properties.
- be able to use the periodic table to locate the positions of elements.

After working through this Factsheet you will:-
- understand the format of the PeriodicTtable.
- understand the term periodicity.
- understand trends in melting point, boiling, electrical conductivity, ionisation energy, atomic radius and electronegativity across the third period.

Exam Hint: *Examination questions across the syllabus will require you to understand the format of the periodic table. Specific questions are often set on describing and explaining the trends in physical and chemical properties down groups and across the third period. Understanding these trends, and the reasons behind them, will also make much of the rest of Chemistry easier to remember.*

The Periodic Table

In the periodic table, elements are placed in order of increasing **atomic number** (fig 1.)

The horizontal rows of elements are called ***periods***

The vertical columns of elements are called ***groups***

Modern periodic law states the properties of elements are a periodic function of their atomic numbers. This means that their properties repeat regularly, so that elements in the same group tend to show similar chemical and physical properties.

You need to know the names of the following groups:

Group Number	Group Name
I	Alkali Metals
II	Alkaline Earth Metals
VII	Halogens
0	Noble Gases

Fig 1. The Periodic Table

Key

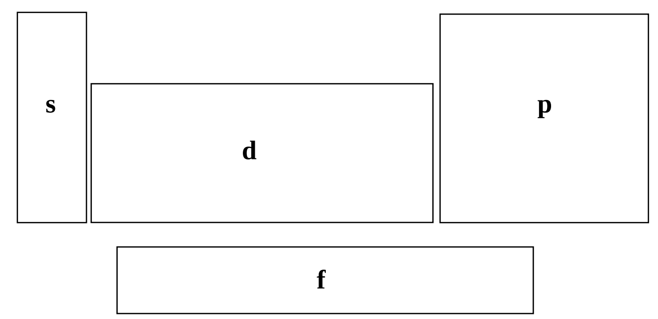
Atomic Number
Symbol
Relative Atomic Mass

	1	2											3	4	5	6	7	0
1	1 H 1																	2 He 4
2	3 Li 7	4 Be 9											5 B 11	6 C 12	7 N 14	8 O 16	9 F 19	10 Ne 20
3	11 Na 23	12 Mg 24											13 Al 27	14 Si 28	15 P 31	16 S 32	17 Cl 35.5	18 Ar 40
4	19 K 39	20 Ca 40	21 Sc 45	22 Ti 48	23 V 51	24 Cr 52	25 Mn 55	26 Fe 56	27 Co 59	28 Ni 59	29 Cu 63.5	30 Zn 65.4	31 Ga 70	32 Ge 73	33 As 75	34 Se 79	35 Br 80	36 Kr 84
5	37 Rb 85	38 Sr 88	39 Y 89	40 Zr 91	41 Nb 93	42 Mo 96	43 Tc (99)	44 Ru 101	45 Rh 103	46 Pd 106	47 Ag 108	48 Cd 112	49 In 115	50 Sn 119	51 Sb 122	52 Te 128	53 I 127	54 Xe 131
6	55 Cs 133	56 Ba 137	57 La 139	72 Hf 178	73 Ta 181	74 W 184	75 Re 186	76 Os 190	77 Ir 192	78 Pt 195	79 Au 197	80 Hg 201	81 Tl 204	82 Pb 207	83 Bi 209	84 Po (210)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Unq (261)	105 Unp (262)	106 Unh (263)												

Lanthanides	58 Ce 140	59 Pr 141	60 Nd 144	61 Pm (147)	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 163	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
Actinides	90 Th 232	91 Pa (231)	92 U 238	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (245)	98 Cf (251)	99 Es (254)	100 Fm (253)	101 Md (256)	102 No (254)	103 Lr (257)

The periodic table is divided into four blocks -s, p, d and f, which tell you the electron sub-shell being filled in that block (see Factsheet 1 - Atomic Structure).

Fig 2. s, p, d and f blocks



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. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

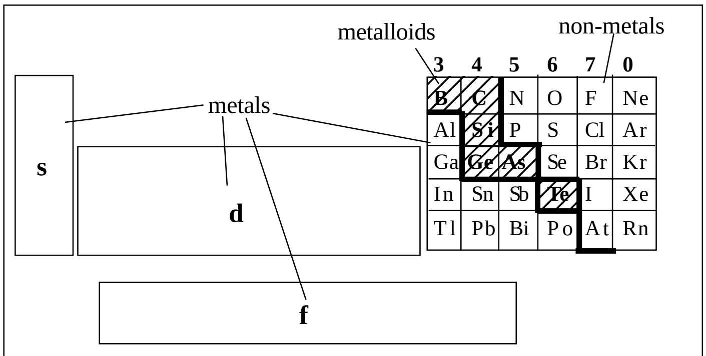
f-block elements: - a block of elements within the transition metals, so called because electrons are being added into the f-subshell in these elements. e.g. Cerium (Ce) Atomic no. 58.
Elec. config. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4s^6 5s^2 4d^{10} 5p^6 6s^2 4f^2$

Metals, Non-metals and Metalloids.
The majority of elements are metals; these are on the left hand side of the periodic table. All of the s, d and f block are metals.

The non-metals are found in the top-right corner of the periodic table; they are all in the p-block.

Along the border between the metals and non-metals there are several elements which are difficult to place, so the name **metalloid** or **semi-metal** is used. Fig 3 shows these elements.

Fig 3. Metals, metalloids and non-metals



To categorise elements, the following criteria are used:-

Metals	• *Good conductors of electricity.* • *Form basic oxides.*
Metalloids	• *Poor conductors of electricity* • *Form amphoteric oxides*
Non-Metals	• *Virtually non-conductors of electricity (insulators)* • *Form acidic oxides.*

Periodic Properties
Elements in the third period (Na, Mg, Al, Si, P, S, Cl, Ar) illustrate important trends in properties across the periodic table. Table 1 overleaf summarises these. When looking at trends in these properties, it is important to be able to explain them using your knowledge of structures.

Melting and Boiling Points
Rise from Na to Si, then fall to low values.

Strong **metallic bonding** in Na, Mg and Al cause high melting and boiling points. The strength of the metallic bonding depends on the **number of outer electrons** - the greater the number, the stronger the bond. So the strength of the metallic bond increases from Na to Al, as do the melting points and boiling points.

Si has a giant covalent structure, leading to very high melting and boiling points.

P, S and Cl have simple molecular structures, held together by Van der Waals forces only, so melting and boiling points are low.

Ar exists as atoms, not molecules. This means that the Van der Waals forces are very weak, giving it a very low melting and boiling point.

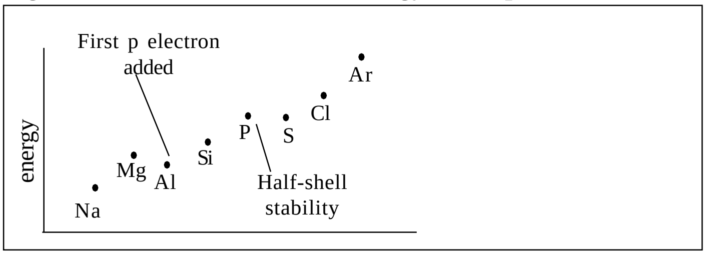
Electrical Conductivity
Relatively high for metals - Na, Mg and Al (due to the de-localised electrons in metallic structures). Lower in metalloids (Si) and almost negligable for non –metals (P, S, Cl and Ar).

First Ionisation Energy
The general trend is for an increase across a period, due to the increasing nuclear change making it more difficult to remove an electron.

Some anomalies occur; these are due to:
- full sub-shell stability - it is comparatively harder to remove an electron from a full sub-shell
- half-shell stability - it is comparatively harder to remove an electron from a half-empty p sub-shell
- p-electrons being easier to remove than s-electrons, due to being at a higher energy level.

This produces the pattern shown below (fig 4) (see Factsheet 1 - Atomic Structure for more details):

Fig 4. Trends in first ionisation energy across period 3



Atomic Radius
Decreases across the period

If we move from one element to the next across the third period, electrons are being added to the same third shell at about the same distance from the nucleus, whilst protons are being added to the nucleus. The increased nuclear change pulls the electrons in the third shell closer to it, hence atomic radii decrease.

Electronegativity
This is the ability to attract electrons within a bond (see Factsheet 5 - Bonding). As we move across the period, electronegativity increases to Cl. Ar has little or no electronegativity since as a noble gas, it has only completed shells of electrons and so does not form bonds at all readily.

Table 1. Physical properties across Period 3

Group	1	2	3	4	5	6	7	0
Element	Sodium	Magnesium	Aluminium	Silicon	Phosphorus	Sulphur	Chlorine	Argon
Symbol	Na	Mg	Al	Si	P	S	Cl	Ar
Character	metallic	metallic	metallic	metalloid	non-metallic	non-metallic	non-metallic	non-metal
Structure	giant metallic	giant metallic	giant metallic	giant covalent	molecular	molecular	molecular	atomic
Melting Point (K)	371	922	933	1683	37	392	172	84
Boiling Point (K)	1156	1380	2740	2628	553	718	238	87
Conductance × 1000 ($ohm^{-1}cm^{-4}$)	10	16	38	4	10^{-16}	10^{-22}	-	-
1st Ionisation Energy ($kJmol^{-1}$)	496	738	578	789	1012	1000	1215	1521
Atomic covalent radius (nm)	0.156	0.136	0.125	0.117	0.110	0.104	0.099	0.095

Practice Questions

1. The *Periodic Table* arranges the different elements in a pattern according to the structure of their atoms and the way in which they behave.

   a) (i) In what order are the elements arranged in the Periodic Table?
   (ii) What name is given to the vertical columns?
   (iii) What name is given to the horizontal rows?

   b) What names are given to the following groups in the Periodic Table?
   (i) Group 1 (ii) Group 2 (iii) Group 7 (iv) Group 0

   c) What name is given to the block of elements found between Groups 2 and 3?

   d) What is the connection between the electronic configuration of an element and:
   (i) the group in which the element is found?
   (ii) the period in which the element is found?

   e) *Periodicity* is the study of the patterns of properties of the elements found in the Periodic Table. State the *general* trends in the variation of the following properties of the elements Na → Ar across Period 3. In each case state the major factor which influences this change.
   (i) Atomic radius (ii) Melting point (iii) First ionisation energy

2. a) (i) Define the *electronegativity* of an element.
   (ii) State and explain how the electronegativities of elements vary across a period of the Periodic Table.

   b) State how the elements of Period 3 (Na – Ar) vary with regard to their tendency to undergo ionic bonding and covalent bonding.

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

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

   a) (i) State and describe the type of bonding present in Na, Mg and Al.
   (ii) Explain why the electrical conductivity of these elements is relatively high.
   (iii) Why does electrical conductivity increase from Na to Mg to Al?

Answers

1. a) (i) Increasing atomic number / number of protons
   (ii) Groups (iii) Periods
   b) (i) Alkali metals (ii) Alkaline earth metals (iii) Halogens
   (iv) Noble gases / inert gases
   c) Transition metals / d-block elements
   d) (i) Group number indicates the number of electrons found in an atom's outer shell (1)
   (ii) In a period, atoms of all elements have the same electron core
   e) (i) *Trend* **Decreases** along the period (1)
   Cause Increasing nuclear charge / number of protons in the nucleus
   (ii) *Trends* Increase Na → Si (1)
   Decrease P → Ar (1)
   Causes Na → Si stronger bonding (1)
   P → Ar weaker van der Waals' forces (1)
   (iii) *Trend* **Increases** along the period (1)
   Cause Increasing nuclear charge (1)

2. a) (i) A measure of the power of an atom of that element, within a molecule (1) to attract towards itself (1) the electrons of a covalent bond (1)
   (ii) Electronegativities **increase** (1)because the tendency of atoms to gain electrons increases (1) (*Or* electron affinities increase (1)) due to an increase in nuclear charge (1) and decreased atomic radius (½) with no increase in shielding (½) Simultaneously, the tendency of atoms to lose electrons decreases (1)
   (*Or* ionisation energies increase (1))

3. a) (i) *Type of bonding* Metallic bonding (1)
   Description Lattice of **cations** (1) with delocalised valency electrons / flux (or 'sea') of electrons (1)
   (ii) Electrons (1) can **move** under an applied p.d. (1)throughout the entire metal (1)
   (iii) Number of outer shell electrons increases from 1 to 2 to 3 (1)
   Hence there are more free / delocalised / mobile electrons (1)

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