Chem Factsheet

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Groups 1 & 2

To succeed with this topic you need to:
- **Understand terms describing physical properties of chemicals (e.g. melting point, density etc)**
- **Understand the use of chemical formulae in balanced equations (Factsheet 03 Moles and Equations)**
- **Understand atomic structure, electronic configuration and the concept of atomic and ionic radii (Factsheet 01 Atomic Structure)**

After working through this Factsheet you will be able to:
- **Recall characteristic physical properties of group 1 and 2 elements**
- **Understand trends in ionisation energies and reactivity within the groups**
- **Recall relevant chemical reactions of the elements**
- **Recall characteristic flame colours**
- **Recall trends in solubilities of sulphates and hydroxides of groups 1 and 2**
- **Recall trends in thermal stabilities of nitrates and carbonates of groups 1 and 2.**

Exam Hint *- Candidates will need to understand trends in the elements and compounds of groups 1 and 2. Importantly, you must be able to explain these trends* ***concisely****. Any explanation must be backed up with evidence, so equations for chemical reactions must be learnt.*

Characteristic physical properties

Group 1

Element	Symbol	Melting Point °C	Boiling Point °C	Density gcm^{-3}	Atomic radius nm	Ionic radius nm
Lithium	Li	181	1330	0.53	0.15	0.06
Sodium	Na	98	890	0.97	0.19	0.10
Potassium	K	63	774	0.86	0.23	0.13
Rubidium	Rb	39	688	1.53	0.24	0.15
Caesium	Cs	29	690	1.88	0.26	0.17

Group 2

Element	Symbol	Melting Point °C	Boiling Point °C	Density gcm^{-3}	Atomic radius nm	Ionic radius nm
Beryllium	Be	1278	2477	1.85	0.11	0.03
Magnesium	Mg	649	1110	1.74	0.16	0.07
Calcium	Ca	839	1487	1.54	0.20	0.10
Strontium	Sr	769	1380	2.60	0.21	0.11
Barium	Ba	725	1640	3.57	0.22	0.13

Members of groups 1 and 2 are all metals. They are silver in colour and tarnish in air. Many of the characteristic trends in their physical properties can be explained by the fact that they show weak metallic bonding (See Factsheet 5).

Group 1 metals have only 1 outer valence shell electron per atom – so only 1 electron per atom is donated into the delocalised electron system.

Group 2 metals have 2 outer valence shell electrons and so have slightly stronger metallic bonding – this is represented in their physical properties.

Exam Hint *- Examiners may require you to know the trends in the physical properties and the explanation behind them.*

Melting Points & Boiling Points
- As group 1 or 2 is descended, melting and boiling points decrease, as the strength of metallic bonding in the elements decreases.
- As atoms/ions get larger, the distance between the delocalized electrons and nucleii increases, so metallic bonding becomes weaker.
- Notice melting and boiling points are higher in group 2, due to stronger metallic bonding as 2 electrons per atom are donated into the delocalised system.

Density
- In general the density rises as group 1 or 2 is descended, due to the fact that the mass of the atoms increases more rapidly than their size with increasing atomic number.
- Notice Group 2 metals are denser than group 1 – due to closer packing of ions, due to stronger bonding.

Atomic/ionic radii
- Increase down groups as number of shells of electrons increase

Comparison with other metals
The physical properties of the group 1 and 2 elements differ from those of most metals in several ways
- They are soft and can be cut with a knife
- Their melting and boiling temperatures are low
- They have low densities (Li, Na and K are less dense than water, and so float on it)

Flame Colours
The electrons within the atoms/ions of group 1 or 2 metals can be excited to higher energy levels (Factsheet 01 – Atomic Structure) by heat energy. When these excited electrons return to their lower, ground state, energy levels, they emit energy – often with a wavelength in the visible light region - hence each group 1 or 2 metal burns with a characteristic flame colour. These characteristic flame colours are also shown by the compounds of these elements – hence they can be used in chemical analysis (i.e. identification of salts)

Flame Colours

Group 1	**Group 2**
Li – carmine red	Mg – white
Na – yellow	Ca – brick red
K – lilac	Sr – crimson red
	Ba – apple green

Ionisation Energies
In a large atom (e.g. Cs) there is:

Group 1

Li
Na
K
Rb
Cs $\downarrow$ As group 1 is descended, ionisation energies **decrease**

- Increased distance between nucleus and outer shell electron
- Increased "shielding effect" as more electrons exist between nucleus and outer electron

Hence, the larger the atom, the easier it is to remove electrons.

Note: Second ionisation energy will be much greater than first, as electrons from a different, lower energy level would be removed.
Again, increasing atomic size leads to decreasing ionisation energies.

Group 2

Be
Mg
Ca
Sr
Ba $\downarrow$ As group 2 is descended, ionisation energies **decrease**

The ionisation energies of group 2 are higher than for group 1, as nuclear charge is greater. Compare sodium and magnesium – sodium has a nuclear charge of +11, whereas magnesium has a nuclear charge of +12. Magnesium is a slightly smaller atom than sodium, hence the attraction for the outer electrons in magnesium is larger, so they are more difficult to remove.

Note: In group 2 there is only a slight increase in ionisation energies from first IE to second IE, but there is a large jump to the third IE as this would involve removing electrons from a lower energy level.

Chemical properties and reactions of groups 1 and 2

Li
Na
K
Rb
Cs $\downarrow$ Reactivity **increases**

Be
Mg
Ca
Sr
Ba $\downarrow$ Reactivity **increases**

The lower the ionisation energies, the easier to remove electrons are, and hence the more reactive the metals are. Hence reactivity increases down both groups, and group 1 is more reactive than group 2.

Group 1 and 2 metals form ionic compounds (with the exception of beryllium) and the only oxidation states shown are:

Group 1: +1
Group 2: +2

Group 1 metals with oxygen
Group 1 metals react with oxygen vigorously. All except Li form a variety of oxides:

Simple oxides containing the O^{2-} ion (this is the only oxide formed by Li)

e.g. $4Na(s) + O_2(g) \rightarrow Na_2O(s)$

Peroxides containing the O_2^{2-} ion (formed in excess oxygen by all of group 1 except Li)

e.g. $2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$

Superoxides containg the radical anion O_2^- (formed by K, Rb, Cs in excess oxygen). NB: these are coloured compounds

e.g. $K(s) + O_2(g) \rightarrow KO_2(s)$

Exam Hint *- You will be expected to know the equations for the formation of oxides for Group1*

Group 2 metals with oxygen
Vigorous reactions forming simple metal oxides $M^{2+}\ O^{2-}$

e.g. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Reactions of Groups 1 and 2 metals with water

Group 1 and most Group 2 metals react with water to produce metal hydroxides and hydrogen gas.

e.g. Group 1: $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Group 2: $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

Reactivity increases as each group is descended, and group 1 metals react more vigorously than group 2

Note: Group 1 metals are stored in oil to stop them reacting with water vapour in the air.

Note: Beryllium **does not react** with water. Magnesium only reacts **very slowly** with cold water, but reacts rapidly with **steam** to form the **oxide** and hydrogen

e.g. $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$

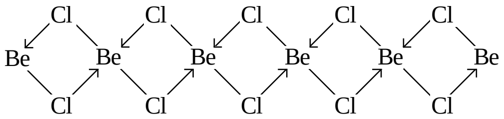
Reactions of Groups 1 and 2 with chlorine gas

Groups 1 and 2 react when heated in chlorine gas to give the chloride:

e.g. Group 1 $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$

Group 2 $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

All group 1 and 2 chlorides are ionic, except for $BeCl_2$ which is covalent – it forms a polymer, using dative bonds to link the linear $BeCl_2$ molecules



Now we shall look at some of the reactions of the oxides of the Group 1 and 2 metals

Reactions of Group 1 and 2 oxides with water

All of the simple metal oxides (containing the O^{2-} anion) react with water to form the metal hydroxide

Ionic equation: $O^{2-}(s) + H_2O(l) \rightarrow 2OH^-(aq)$

e.g. Group 1 $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

Group 2 $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

Note: MgO is insoluble in water, so will not react in this fashion

Peroxides (containing the O_2^{2-} anion) react with water to form the metal hydroxide and hydrogen peroxide:

Ionic equation: $O_2^{2-} + 2H_2O(l) \rightarrow 2OH^-(aq) + H_2O_2(aq)$

e.g. $Na_2O_2(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2O_2(aq)$

Superoxides (containing the O_2^- anion) react with water to form the metal hydroxide, hydrogen peroxide and oxygen gas.

Ionic equation: $2O_2^-(s) + 2H_2O(l) \rightarrow 2OH^- + H_2O_2(aq) + O_2(g)$

e.g. $KO_2(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2O_2(aq) + O_2(g)$

Note: all solutions formed will be alkaline, as the OH^- ion is produced in each case.

The reactions are of the anion, so the metal cation is irrelevant (unless Mg^{2+}, as MgO is insoluble)

Reactions of Group 1 and 2 oxides with acids

The metal oxides are basic compounds, so they will all react will acids to form a salt and water

e.g. Group 1 $Na_2O(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

Group 2 $MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$

Solubilities of Group 2 hydroxides

$Mg(OH)_2$ — insoluble
$Ca(OH)_2$
$Sr(OH)_2$
$Ba(OH)_2$ — soluble

As the group is descended, solubility **increases**

Solubilities of Group 2 sulphates

$MgSO_4$ — soluble
$CaSO_4$
$SrSO_4$
$BaSO_4$ — insoluble

As the group is descended, solubility **decreases**

> ***Exam Hint*** *- You need to learn the above solubility trends, but do not need to be able to explain them at AS level*

Trends in thermal stabilities of nitrates and carbonates

Thermal decomposition is more likely in those compounds if the metal cation polarises the anion (CO_3^{2-} or NO_3^-).
Consequently:

- Group 2 compounds (+2 charge on cation) are more likely to decompose than group 1 compounds (+1 charge on cation)
- The smaller the cation (near the top of the group) the more likely they are to decompose.

Nitrates

Group 1 nitrates (except lithium nitrate) decompose to give the metal nitrite and oxygen gas

e.g. $2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$

Group 2 nitrates (and lithium nitrate) decompose to give the metal oxide, the brown gas nitrogen dioxide and oxygen:

e.g. $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$

$4LiNO_3(s) \rightarrow 2Li_2O(s) + 4NO_2(g) + O_2(g)$

Thermal stability increases down both groups

Carbonates

Group 1 carbonates will not decompose on heating, except lithium carbonate:

$Li_2CO_3(s) \rightarrow Li_2O(s) + CO_2(g)$

Group 2 carbonates all decompose (except for barium carbonate, which is stable) to form the metal oxide and carbon dioxide gas

e.g. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Thermal stability increases down both groups

Practice Questions

1. Explain why both group 1 and 2 elements have low melting points for metals, and why group 1 melting points are lower than those of group 2.

2. Describe a test which would distinguish sodium chloride from potassium chloride.

3. Explain the trend in first ionisation energies as group 1 is descended.

4. Write balanced equations for the following reactions:
   (a) strontium with oxygen

   (b) potassium with water

   (c) magnesium with steam

   (d) sodium with chlorine

5. Which is the only group 2 metal not to form ionic compounds?

6. State which of each pair of compounds is more soluble:
   (a) magnesium hydroxide or barium hydroxide

   (b) calcium sulphate or strontium sulphate

7. Give two reasons as to why rubidium nitrate is more thermally stable than magnesium nitrate

8. Write balanced equations for the following thermal decompositions:
   (a) sodium nitrate

   (b) magnesium nitrate

   (c) magnesium carbonate

Answers

1. Group 1 and 2 metals both have relatively weak metallic bonding. Metallic bonding in group 1 is weaker than that in group 2, as group 1 metals only donate 1 electron per atom into the delocalised sea of electrons, whereas group 2 metals donate 2 outer valence shell electrons per atom.

2. Flame test: sodium chloride gives a golden yellow flame, and potassium chloride gives a lilac flame.

3. As group 1 is descended, first ionisation energies decrease, due to increased distance between the nucleus and outer electron and increased shielding.

4. (a) $2Sr(s) + O_2(g) \rightarrow 2SrO(s)$

   (b) $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

   (c) $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$

   (d) $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

5. Beryllium

6. (a) barium hydroxide

   (b) calcium sulphate

7. larger cation, so less polarising power
   lower cation charge, so less polarising power

8. (a) $2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$

   (b) $2Mg(NO_3)_2 \rightarrow 2MgO(s) + 4NO_2\ (g) + O_2(g)$

   (c) $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

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