Trends down a group: Group 1 – The alkali metals

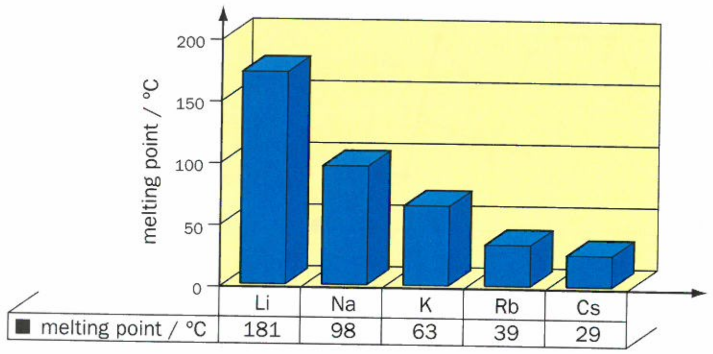
The metals in this first group don't have many uses as the elements themselves. They are too reactive.
However, you will certainly use some of their compounds every day.
Here are their electronic configurations:

	1	
lithium	Li	$1s^2 2s^1$
sodium	Na	$1s^2 2s^2 2p^6 3s^1$
potassium	K	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
rubidium	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$
caesium	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$

Physical properties

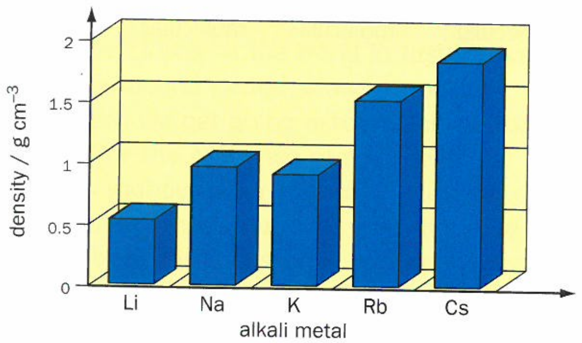
Can you remember why sodium, and the other **alkali metals**, are unusual metals?
They have ***low melting points*** and are ***very soft***.
They can be cut with a knife. Look at the photo opposite:

The graph below shows how their melting points change with increasing atomic number:



- Can you see a trend going down the group?
- What can you say about the melting points of the alkali metals going down the group?

For metals, the elements in Group 1 also have ***very low densities***.
Look at their densities on the graph below:



- What is the *general* trend going down the group?
- Which elements break the pattern?

Why are the alkali metals stored under paraffin (oil)?

Sodium is a soft metal

The alkali metals can be cut with a knife

Chemical properties

The alkali metals are the most reactive group of metals in the Periodic Table. Look at the photo of caesium reacting with water:

Lithium fizzes on the surface of the water, giving off hydrogen gas.
Sodium gets hot enough to melt itself as it reacts.
Potassium gets so hot that it ignites the hydrogen gas given off.
It burns with a lilac flame (the colour is caused by K^+ ions).
Can you see a pattern in reactivity going down the group?

The alkali metals get *more reactive* going down the group.

The Periodic Table is very useful. Its groups make chemistry easier!
You only have to learn the reactions of one element in a group.
The others are usually similar. For example,

lithium + water ⟶ lithium hydroxide (alkaline solution) + hydrogen

$$2\,Li(s) + 2\,H_2O(l) \longrightarrow 2\,LiOH(aq) + H_2(g)$$

Knowing this, we know the equations for the other alkali metals:

sodium + water ⟶ sodium hydroxide + hydrogen

$$2\,Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

etc.

We can explain the patterns in reactivity by looking at the atomic structure of the elements.

Why are the alkali metals (from Group 1) so reactive?
Like many atoms, they react to get the electronic configuration of a noble gas. As the elements, they have just one electron in their atom's outer shell.
So when they react, they lose that one outer electron, leaving a complete octet (s^2p^6) shell. (See ionic bonding, Chapter 4.)
Look for the alkali metals on the graph of first ionisation energies on page 100: Group 1 elements are so reactive because it is so easy for them to lose just one electron.

- Why do you think that sodium, from Group 1, is more reactive than its neighbour from Group 2, magnesium?

Why do the metals get more reactive descending Group 1?
The outer electron gets easier to remove as you go down the group.
Remember that electrons are negative. They are attracted to the positive protons in the nucleus. As you go down the group, the atoms get bigger.
Therefore the outer electron gets ***further away from the attractive force of the nucleus***.
There are also more inner shells full of electrons between the outer electron and the nucleus. This has a **shielding effect** on the outer electron, again reducing the attraction between it and the nucleus.

These two factors outweigh the increasing nuclear charge as the atomic (proton) number increases, making it easier for an electron to be removed from a larger atom.

Caesium reacts explosively with water!

The Group 1 metals form alkaline solutions (soluble hydroxides) with water