

1		2		Group										13	14	15	16	17	18
				H										←metals non-metals→					He
Li	Be											B	C	N	O	F	Ne		
Na	Mg											Al	Si	P	S	Cl	Ar		
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr		
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe		
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn		
Fr	Ra																		

Melting Points:

Melting points increase from Group 1 to Group 14 due to a change in bonding from metallic (Groups 1, 2 and 13) to giant covalent structures (Group 14). From Group 14 to Group 18, melting points decrease as the nature of the bonding within the elements changes from giant covalent structures (Group 14) to simple covalent molecules (Groups 15, 16 and 17) and eventually monatomic atoms (Group 18).

Electrical Conductivity:

This decreases across a Period from the left-hand-side to the right-hand side as the nature of the elements changes from metallic (good conductors of electricity due to their delocalised electrons) to metalloid (semiconductors) to non-metallic (insulators).

Acid / Base Nature of Oxides:

The oxides of the metallic elements are basic. Those of the metalloids tend to be amphoteric while those of the non-metallic elements are acidic.

Atomic Radius:

The atomic radius of an atom of an element decreases across a Period from the left-hand-side to the right-hand-side, *i.e.* within the Third Period, sodium has the largest atomic radius while argon has the smallest. This is due to the increasing effective nuclear charge, *i.e.* sodium has 11 protons in its nucleus while argon has 18. The additional protons in the nucleus of argon exert a stronger electrostatic force of attraction on the valence electrons, pulling them closer to the nucleus.

Oxidising and Reducing Power:

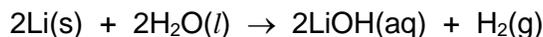
The metallic elements on the left-hand-side of the Periodic Table are good reducing agents because they readily lose electrons to other chemicals. The non-metallic elements on the right-hand-side of the Periodic Table are good oxidising agents because they readily remove electrons from other chemicals.

The Group 1 Elements (Alkali Metals)

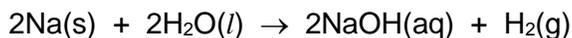
The Group 1 elements are a collection of soft, highly reactive metals that have low melting points and low densities. They all have a single electron in their valence shell, which is lost when they react, resulting in the formation of a cation with a single positive charge.

The Group 1 metals all react vigorously with cold water to form an alkaline solution of the metal hydroxide and hydrogen gas. Lithium, sodium and potassium all float on the surface of the water. The hydrogen gas produced by the sodium ignites and burns with an orange flame while the hydrogen gas produced by the potassium burns with a lilac flame:

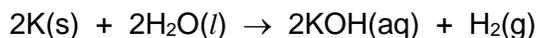
Lithium + water → lithium hydroxide + hydrogen



sodium + water → sodium hydroxide + hydrogen

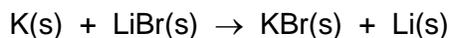


potassium + water → potassium hydroxide + hydrogen



Reactivity increases upon descending the Group, lithium being the least reactive while caesium is the most reactive. Reactivity increases due to a decrease in the electrostatic force of attraction between the nucleus and single valence electron. This decrease in the electrostatic force of attraction is due to an increasing atomic radius, and increased shielding effect that is observed while descending the Group.

A more reactive Group 1 metal can displace a less reactive Group 1 metal from its compounds. For example, potassium is more reactive than lithium. Therefore, potassium can displace lithium from the compound lithium bromide:



The Group 17 Elements (Halogens)

The Group 17 elements are all diatomic non-metals which vary widely in their appearance:

Fluorine, F_2 , is a pale yellow gas.

Chlorine, Cl_2 , is a yellow-green gas.

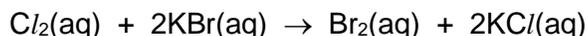
Bromine, Br_2 , is a volatile reddish-brown (dark orange) liquid.

Iodine, I_2 , is a blackish-purple solid.

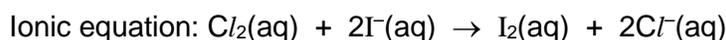
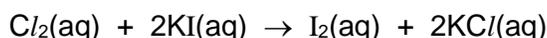
They all have seven valence electrons and react by attracting a single electron into their valence shell to form either an anion with a single negative charge (if reacting with a metal) or a single covalent bond (if reacting with another non-metal).

Unlike the Group 1 metals, reactivity decreases while descending the Group, *i.e.* fluorine is the most reactive Group 17 element while iodine is the least reactive. The more reactive elements from the top of the Group can displace the less reactive elements from their compounds, for example:

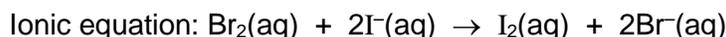
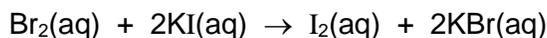
chlorine + potassium bromide → bromine + potassium chloride



chlorine + potassium iodide → iodine + potassium chloride



bromine + potassium iodide → iodine + potassium bromide



In each case, the bromine and iodine that are produced in aqueous solution will dissolve preferentially in a non-polar solvent such as hexane to produce reddish-brown (or orange) and purple solutions respectively.

The Transition Metals

The transition metals are found in a block towards the centre of the Periodic Table. They are typical metals, having high tensile strength, high melting points and high densities. The main properties of the transition metals are that they form coloured compounds, have variable oxidation states and function as catalysts, *i.e.* they increase the rate of chemical reactions.

Iron(II) salts, e.g. FeSO_4 , are pale green.

Iron(III) salts, e.g. FeCl_3 , are yellow.

Copper(II) salts, e.g. CuSO_4 , are blue.

The oxidising agent potassium manganate(VII), KMnO_4 , is purple.

The oxidising agent potassium dichromate(VI), $\text{K}_2\text{Cr}_2\text{O}_7$, is orange.

Reactions of the transition metals and their compounds are often accompanied by colour changes. For example, if an iron(II) salt is oxidised to iron(III), we would expect to see a colour change from pale green to yellow. If copper(II) was displaced from its salt by iron, we would expect to see a colour change from blue to pale green or yellow.