3.2 Properties of elements in group 1 and group 17

Learning objectives

- Understand that elements in the same group have similar chemical properties and show a gradual variation in physical properties
- Describe some reactions of elements in group 1 and group 17
- Describe the changes from basic to acidic oxides across a period
- Write equations for the reactions of oxides with water and predict the acidity of the resulting solutions

Introduction

The chemical properties of an element are determined by the electron configuration of its atoms. Elements of the same group have similar chemical properties as they have the same number of valence electrons in their outer energy level.

Group 18: the noble gases

Group 18, contains the least reactive elements – the noble gases and were discovered at the end of the 19th century after Mendeleyev first published his table.

- They are colourless gases.
- They are monatomic: they exist as single atoms.
- They are very unreactive.

With the exception of helium, they have complete valence energy levels with eight electrons; a **stable octet**. Helium has a complete principal first energy level with two electrons. Their lack of reactivity can be explained by the inability of their atoms to lose or gain electrons.
Other elements loose or gain electrons so as to achieve the electron configuration of their nearest noble gas.

Group 1: the alkali metals

All the elements are silvery metals and are too reactive to be found in nature. They are usually stored in oil to prevent contact with air and water.
The bonding in all these elements is metallic. The solid is held together by electrostatic attraction between the positive ions in the lattice and the delocalised electrons.

Melting point decreases down group 1.

As the ions get larger as we go down the group, the nucleus becomes further from the delocalised electrons and the attraction becomes weaker.

Physical properties	Chemical properties
They are good conductors of electricity and heat. They have low densities. They have grey shiny surfaces when freshly cut with a knife.	• They are very reactive metals. • They form ionic compounds with non- metals.

They form single charged ions, M+, with the stable octet of the noble gases when they react. Their low ionization energies give an indication of the ease with which the outer electron is lost. Reactivity increases down the group as the elements with higher atomic number have the lowest ionization energies.

Reactions of the elements in group 1

The alkali metals react with water to produce hydrogen and the metal hydroxide.

- Lithium floats and reacts slowly. It releases hydrogen but keeps its shape.
- Sodium reacts with a vigorous release of hydrogen. The heat produced is sufficient to melt the unreacted metal, which forms a small ball that moves around on the water surface.
- Potassium reacts even more vigorously to produce sufficient heat to ignite the hydrogen produced. It produces a lilac coloured flame and moves excitedly on the water surface.

The metals are called alkali metals because the resulting solution is alkaline owing to the presence of the hydroxide ion formed.
The reactions become more vigorous going down the group because the ionisation energy decreases as the size of the atom increases.

The general equation for the reaction is:

$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$
Example:
For potassium,
$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

Reactions with halogens

The alkali metals react vigorously with oxygen and all tarnish rapidly in air. The general equation for the reaction is:
$4M(s) + O_2(g) \rightarrow 2M_2O(s)$

M_2O is a basic oxide that will dissolve in water to form an alkaline solution, containing M^+ and OH^- ions.

Group 17: the halogens

The elements in group 17 are known as the halogens. They are all non-metals consisting of diatomic molecules (X2).
Properties:

Element	Symbol	Atomic number	Electron configuration	Colour	Melting point / °C	Boiling point / °C	Physical state at room temperature and pressure
fluorine	F	9	$[He]2s^22p^5$	pale yellow	−220	−188	gas
chlorine	Cl	17	$[Ne]3s^23p^5$	yellow–green	−101	−35	gas
bromine	Br	35	$[Ar]3d^{10}4s^24p^5$	deep red liquid, orange vapour	−7	59	liquid
iodine	I	53	$[Kr]4d^{10}5s^25p^5$	grey shiny solid, purple vapour	114	184	solid

The nuclei have a high effective charge, of approximately +7, and so exert a strong pull on any electron from other atoms.
The attraction is greatest for the smallest atom fluorine, which is the most reactive non-metal in the Periodic Table.
Reactivity decreases down the group as the atomic radius increases and the attraction for outer electrons decreases.
The very high reactivity of fluorine can be explained in terms of an exceptionally weak F–F bond and the strength of the bonds it forms with other atoms.

Reaction with Group 1 metals

The halogens all react with the alkali metals to form salts. The general equation is:

$2M(s) + X_2(g) \rightarrow 2MX(s)$

The salts formed are all white/colourless, fairly typical ionic compounds. They contain M+ and X− ions. All alkali metal chlorides, bromides and iodides are soluble in water and form colourless, neutral solutions.

Displacement reactions

The relative reactivity of the elements can also be seen by placing them in direct competition for an extra electron. When chlorine is bubbled through a solution of potassium bromide the solution changes from colourless to orange owing to the production of bromine:

$$2KBr(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + Br_2\,(aq)$$

$$2Br^-(aq) + Cl_2\,(aq) \rightarrow 2Cl^-(aq) + Br_2\,(aq)$$

The chlorine has gained an electron and so forms the chloride ion, Cl^-. The bromide ion loses an electron to form bromine.

Other reactions are:

$$2I^-(aq) + Cl_2(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$$

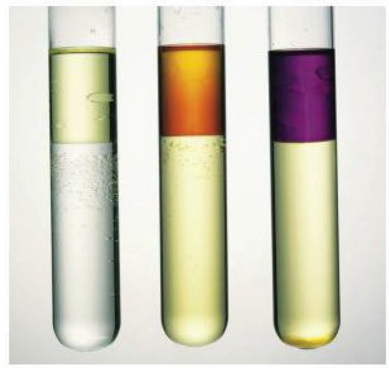
The colour changes from colourless to dark orange/brown owing to the formation of iodine.

$$2I^-(aq) + Br_2(aq) \rightarrow 2Br^-(aq) + I_2(aq)$$

The colour darkens owing to the formation of iodine.

To distinguish between bromine and iodine more effectively, the final solution can be shaken with a hydrocarbon solvent. Iodine forms a violet solution and bromine a dark orange solution as shown in the photo.

The photo shows solutions of chlorine (left), bromine (middle), and iodine (right) in water (lower part) and cyclohexane (upper part).



The halogens form insoluble salts with silver. Adding a solution containing the halide to a solution containing silver ions produces a **precipitate** that is useful in identifying the halide ion.

$$Ag^+(aq) + X^-\,(aq) \rightarrow AgX(s)$$

Summary table for displacement reactions

	KCl(aq)	KBr(aq)	KI(aq)
$Cl_2(aq)$	no reaction	orange solution	dark red–brown solution
$Br_2(aq)$	no reaction	no reaction	dark red–brown solution
$I_2(aq)$	no reaction	no reaction	no reaction

These reactions are all redox reactions (Topic **9**), in which a more reactive halogen oxidises a less reactive halide ion. Chlorine is a stronger oxidising agent than bromine and iodine; it will oxidise bromide ions to bromine, and iodide ions to iodine. Bromine is a stronger oxidising agent than iodine and will oxidise iodide ions to iodine.

Bonding of the Period 3 oxides

Ionic compounds are generally formed between metal and nonmetal elements and so the oxides of elements Na to Al have **giant ionic** structures.

Covalent compounds are formed between non-metals, so the oxides of phosphorus, sulfur, and chlorine are **molecular covalent**. The oxide of silicon, which is a metalloid, has a **giant covalent** structure.

The oxides become more ionic down a group as the electronegativity decreases. The conductivity of the molten oxides gives an experimental measure of their ionic character.

Basic oxides

Oxides of elements may be classified as basic, acidic or amphoteric.

In general, metallic oxides are **basic** and non-metallic oxides are **acidic**.

A basic oxide is one that will react with an acid to form a salt and, if soluble in water, will produce an alkaline solution. Sodium oxide and magnesium oxide dissolve in water to form alkaline solutions owing to the presence of hydroxide ions:

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$
$$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$$

Magnesium oxide, because of the relatively high charges on the ions, is not very soluble in water but it does react to a small extent to form a solution of magnesium hydroxide, which is alkaline

A basic oxide reacts with an acid to form a salt and water. The oxide ion combines with two H+ ions to form water:

$$O2^{-}(s) + 2H^{+}(aq) \rightarrow H_2O(l)$$
$$Li_2O(s) + 2HCl(aq) \rightarrow 2LiCl(aq) + H_2O(l)$$
$$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$$

Amphoteric oxides

Aluminium is on the dividing line between metals and non-metals and forms an amphoteric oxide – these have some of the properties of a basic oxide and some of an acidic oxide. Aluminium oxide does not affect the pH when it is added to water as it is essentially insoluble.

Reaction with acids: $Al_2O_3 + 6H^{+} \rightarrow 2Al^{3+} + 3H_2O$

Example: $Al_2O_3(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2O(l)$

Reaction with alkalis/bases: $Al_2O_3 + 2OH^{-} + 3H_2O \rightarrow 2Al(OH)^{3+}$

	Sodium	Magnesium	Aluminium	Silicon	Phosphor	Sulfur
Formula of oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_2 SO_3
Nature of element	metal			non-metal		
Nature of oxide	basic		amphoteric	acidic		
Reaction with water	soluble, reacts	sparingly soluble, some reaction	insoluble		soluble, reacts	
Solution formed	alkaline	slightly alkaline	–		acidic	

Acidic oxides

The non-metallic oxides react readily with water to produce acidic solutions. Phosphorus (V) oxide reacts with water to produce phosphoric (V) acid:

P_4O_6 (phosphorus (III) oxide) and P_4O_{10} (phosphorus (V) oxide) form phosphoric (III) and phosphoric (V) acid, respectively, when they react with water:

$$P_4O_6(s) + 6H_2O(l) \rightarrow 4H_3PO_3(aq)$$
$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$$

Sulfur trioxide reacts with water to produce sulfuric (VI) acid:

$$SO_3\ (l) + H2O\ (l) \rightarrow H_2SO_4\ (aq)$$

Sulfur dioxide reacts with water to produce sulfuric (IV) acid:

$$SO_2\ (g) + H_2O\ (l) \rightarrow H_2SO_3\ (aq)$$

Dichlorine heptoxide (Cl2O7) reacts with water to produce chloric (VII) acid (HClO4):

$$Cl_2O_7\ (l) + H_2O\ (l) \rightarrow 2HClO_4\ (aq)$$

Dichlorine monoxide (Cl2O) reacts with water to produce chloric (I) acid (HClO):

$$Cl_2O\ (l) + H_2O(l) \rightarrow 2HClO(aq)$$

Silicon dioxide does not react with water, but reacts with concentrated alkalis to form silicates:

$$SiO_2(s) + 2OH–(aq) \rightarrow SiO_3{}^{2-}(aq) + H_2O(l)$$

Nitrogen oxides

There are many oxides of nitrogen, ranging in formula from N2O to N2O5. Two of the most environmentally important are nitrogen (II) oxide (NO) and nitrogen (IV) oxide (NO2).

Nitrogen reacts with oxygen at very high temperatures to form NO.

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

This reaction occurs in the internal combustion engine. NO is virtually insoluble in water and is classified as a neutral oxide.

NO can be oxidised in the atmosphere to NO2, which can react with water to produce nitric(V) acid (HNO3), which is one of the acids responsible for acid deposition.NO2 can be classified as an acidic oxide:

$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

N_2O(nitrogen(I) oxide, nitrous oxide) is another neutral oxide. N_2O is also known as laughing gas and major uses include as an anaesthetic and as the propellant in ‘squirty cream’.