

ESAT

Chemistry

Groups and Metals

Free lesson notes

Covers	C7.1 to C7.3 and C14.1 to C14.5
Format	Lesson, worked examples, self-check
Price	Free

About this lesson

Elements in the same group of the Periodic Table behave in a family way, and metals can be ranked in a league table of reactivity. Together these two ideas let you predict the outcome of reactions you have never seen: will potassium react more violently than sodium? Will bromine push iodine out of a solution? Can you make iron by heating its ore with carbon? Questions of this kind are among the most reliable marks in ESAT Chemistry, because they depend on a handful of patterns and one underlying explanation (how easily an atom loses or gains electrons).

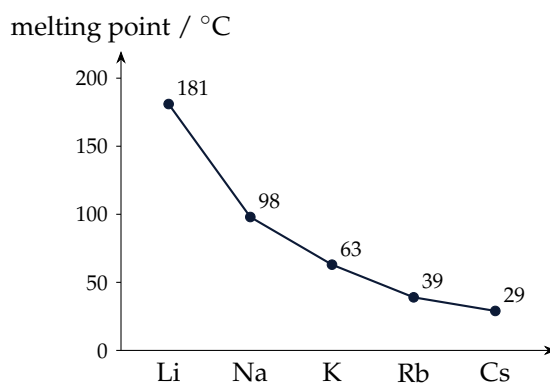
The lesson begins with the three families your specification names: the alkali metals (Group 1), the halogens (Group 17, called Group 7 in older books) and the noble gases (Group 18). We then widen to metals in general: the reactivity series, displacement reactions, how metals are extracted from their ores, the uses of common metals and alloys, and the special features of the transition metals. Specification codes: C7.1–C7.3 (group chemistry) and C14.1–C14.5 (metals). Relative atomic masses are always given in the exam where they are needed, so there is nothing to memorise except the patterns and the facts in the boxes.

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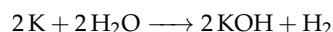
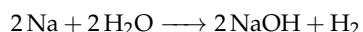
1 Group 1: the alkali metals

The family. Group 1 contains lithium (Li), sodium (Na), potassium (K), rubidium (Rb) and caesium (Cs) (and francium). Each atom has *one electron in its outer shell* (Li 2, 1; Na 2, 8, 1; K 2, 8, 8, 1), which is why they form ions with a charge of 1+ and why they behave so similarly. They are called *alkali* metals because they react with water to give alkaline solutions.

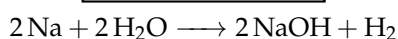
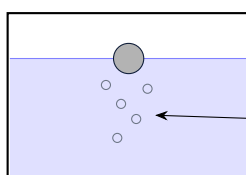
Physical properties. They are *soft* metals (you can cut sodium with a knife), with low densities (lithium, sodium and potassium float on water) and low melting points compared with most metals. The melting point *falls* as you go down the group.



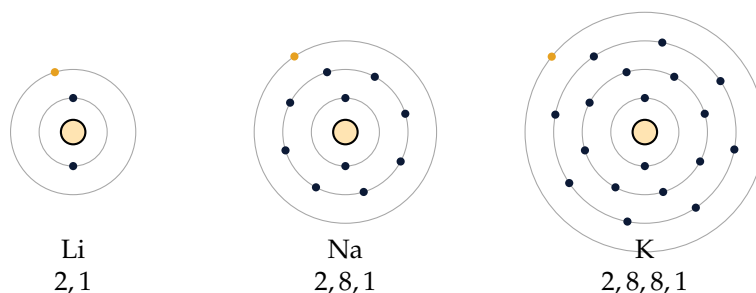
Reaction with water. All the alkali metals react with water, giving the metal hydroxide (an alkaline solution) and hydrogen:



The general pattern is $2\text{M} + 2\text{H}_2\text{O} \longrightarrow 2\text{MOH} + \text{H}_2$ for any Group 1 metal M. Universal indicator turns purple (strongly alkaline). Lithium fizzes steadily; sodium melts into a ball (the reaction releases heat) and darts about; potassium reacts so vigorously that the hydrogen ignites with a lilac flame; rubidium and caesium react explosively.



Reactivity increases down the group. A Group 1 atom reacts by *losing* its one outer electron. Going down the group the outer electron is in a shell further from the nucleus and is more *shielded* from the nuclear attraction by the inner shells, so the attraction on it is weaker and it is lost more easily. Hence potassium is more reactive than sodium, which is more reactive than lithium.



outer electron further from the nucleus and more shielded: lost more easily, reactivity increases

Other reactions and storage. They react with oxygen (tarnishing quickly in air) and with the halogens to form white ionic solids such as NaCl; their compounds are ionic, mostly colourless and mostly soluble. Because they react with air and water the more reactive ones are stored under oil.

Worked example: predicting a reaction

Rubidium (Rb) is below potassium in Group 1. Predict its reaction with water and write the equation. Rubidium is more reactive than potassium, so it reacts even more violently (explosively), giving rubidium hydroxide and hydrogen: $2\text{Rb} + 2\text{H}_2\text{O} \longrightarrow 2\text{RbOH} + \text{H}_2$. The solution is strongly alkaline. Check the equation is balanced: Rb 2 = 2; H 4 = 2 + 2; O 2 = 2.

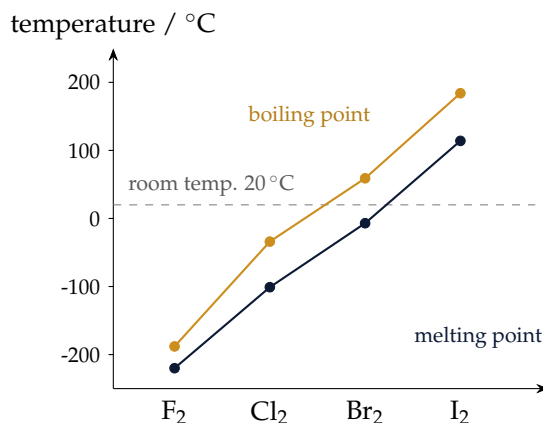
Common mistakes

- Saying the metal gets more reactive down the group because it has more electrons to lose. It still loses only *one*; it is the distance and shielding that make it easier.
- Writing the formula of the hydroxide wrongly: sodium hydroxide is NaOH (one OH^- for each Na^+).
- Thinking the melting point rises down the group as in Group 17. For Group 1 it falls.

2 Group 17: the halogens

The family. Fluorine (F), chlorine (Cl), bromine (Br) and iodine (I) have *seven* electrons in the outer shell and are non-metals. They exist as **diatomic molecules** (F_2 , Cl_2 , Br_2 , I_2) held by a single covalent bond, and each atom reacts by *gaining* one electron to form a halide ion (Cl^- , $1-$) or by sharing one pair.

Trends in physical properties. At room temperature fluorine is a pale yellow gas, chlorine a pale green gas, bromine a red-brown liquid and iodine a grey solid (it gives a purple vapour when heated). The melting and boiling points *increase* down the group, because the molecules get larger and the intermolecular forces between them get stronger (they are simple molecular substances: the covalent bond within each molecule is not broken).



Reactivity decreases down the group. A halogen atom reacts by *gaining* an electron into its outer shell. Going down, the outer shell is further from the nucleus and more shielded, so an incoming electron is attracted less strongly: iodine gains electrons less readily than chlorine. This is the opposite trend to Group 1 (where the outer electron is *lost*).

Displacement reactions. Because of this trend, a more reactive halogen will *displace* (push out) a less reactive halogen from a solution of its salt. Chlorine displaces bromine and iodine; bromine displaces iodine but not chlorine; iodine displaces neither.



The evidence is a colour change: when bromine is formed the colourless bromide solution turns orange, and when iodine is formed it turns brown.

chlorine water
+ potassium bromide



orange: bromine
formed

bromine water
+ potassium iodide



brown: iodine formed

bromine water
+ potassium chloride



no change

In ionic terms, the chlorine atom has gained an electron from the bromide ion: $\text{Cl}_2 + 2 \text{Br}^- \longrightarrow 2 \text{Cl}^- + \text{Br}_2$.

Worked example: will it react?

Which of these mixtures will react? (a) bromine water with sodium chloride solution; (b) chlorine water with sodium iodide solution.

(a) Bromine is less reactive than chlorine, so it cannot displace chlorine: no reaction. (b) Chlorine is more reactive than iodine, so it displaces it: $\text{Cl}_2 + 2 \text{NaI} \longrightarrow 2 \text{NaCl} + \text{I}_2$; the solution turns brown.

Exam tip

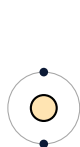
For displacement, write the three halogens in order of reactivity (Cl, Br, I). A halogen only displaces the ones *below* it. The same rule works for metals in the next section: a more reactive element displaces a less reactive one from its compound.

Common mistakes

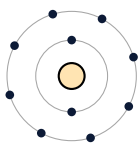
- Predicting that bromine displaces chlorine, because “bromine is bigger”. Reactivity falls down Group 17.
- Writing Cl instead of Cl_2 in equations: the elements are diatomic.
- Forgetting that a displacement needs the halogen *element* and the *halide* ion in solution; a halogen added to its own halide gives nothing.

3 Group 18: the noble gases

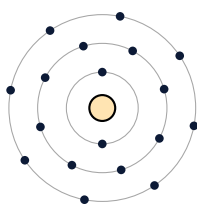
The family. Helium, neon, argon, krypton, xenon and radon have full outer shells (2 for helium, 8 for the rest), so they have no tendency to gain, lose or share electrons. They are **unreactive** (inert), exist as single atoms (*monatomic*), and are colourless gases at room temperature.



He 2



Ne 2, 8



Ar 2, 8, 8

full outer shells:
no need to gain, lose
or share electrons

Trend. The boiling points rise down the group (He -269°C , Ne -246°C , Ar -186°C , Kr -153°C , Xe -108°C) because the atoms get larger, giving stronger intermolecular forces between them. **Uses** depend on being unreactive or very light:

helium (balloons and airships; it is less dense than air and, unlike hydrogen, it is not flammable); argon (filament lamps and as a shield during welding, to stop the hot metal reacting with oxygen); neon (glowing signs).

Worked example: explaining a use

Why is argon, not nitrogen, used inside a filament lamp?

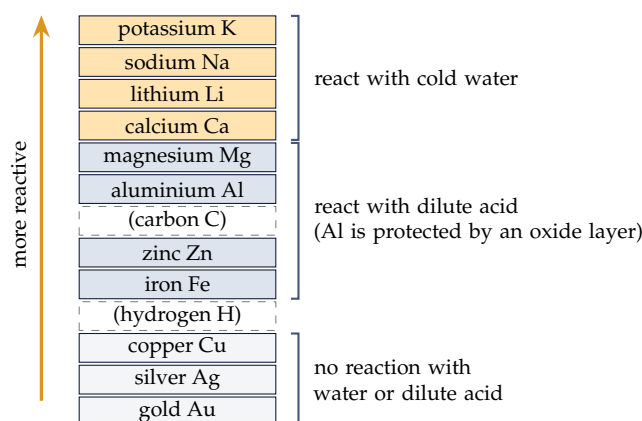
The tungsten filament is white hot. At that temperature it would react with oxygen (and with nitrogen, in some conditions). Argon is unreactive because its outer shell is full, so it protects the filament and the lamp lasts longer.

Common mistakes

- Saying noble gases are unreactive because they have “eight outer electrons”; helium has two (a full first shell).
- Writing the noble gases as diatomic, e.g. Ne_2 : they are single atoms.
- Thinking helium is used in balloons because it is unreactive and nothing more: the key property is its low density.

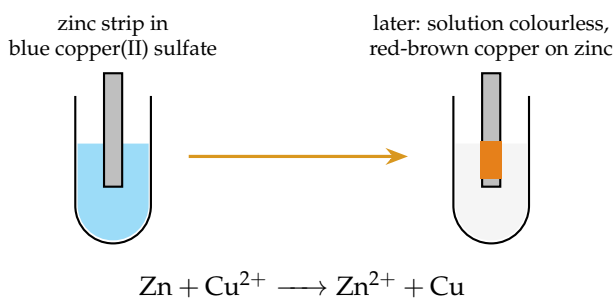
4 The reactivity series of metals

Metals lose electrons. A metal atom reacts by losing its outer electrons to form a positive ion (for example $\text{Mg} \longrightarrow \text{Mg}^{2+} + 2\text{e}^-$). The more easily an atom loses electrons, the more reactive the metal. Experiments with water, acid and oxygen allow us to rank the metals in the **reactivity series**. Hydrogen and carbon are included for reference.



Reactions. With dilute acid a metal above hydrogen gives a salt and hydrogen: $\text{Mg} + 2\text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$. The more reactive the metal, the more vigorous the fizzing (calcium very fast, magnesium fast, zinc steady, iron slowly). Copper, silver and gold are below hydrogen and do not react. With steam, magnesium forms magnesium oxide and hydrogen. Aluminium *appears* unreactive in water and acid because of a thin, tough layer of aluminium oxide on its surface, but it is in fact quite reactive once that layer is removed.

Displacement of metals. A more reactive metal displaces a less reactive metal from a solution of its compound. When zinc is placed in blue copper(II) sulfate solution, zinc atoms give electrons to copper ions:



Zinc has lost electrons (so it is oxidised) and copper ions have gained them. Copper is deposited as a pink-brown solid, the blue colour fades and the mixture warms. Copper, being less reactive than zinc, does *not* displace zinc from zinc sulfate.

Worked example: using displacement to rank metals

Metal A displaces metal B from its sulfate solution, but metal C does not displace metal A from its sulfate solution. Put A, B and C in order of reactivity (most reactive first), given that C is less reactive than B. A displaces B, so A is more reactive than B. C does not displace A, which is consistent with C being less reactive than A. With C less reactive than B: $A > B > C$.

Worked example: a displacement equation

Write the balanced equation for magnesium reacting with iron(II) sulfate solution. Magnesium is above iron, so it displaces it: $\text{Mg} + \text{FeSO}_4 \longrightarrow \text{MgSO}_4 + \text{Fe}$. Mg: 1 = 1; Fe: 1 = 1; SO_4 : 1 = 1. Ionic form: $\text{Mg} + \text{Fe}^{2+} \longrightarrow \text{Mg}^{2+} + \text{Fe}$.

Common mistakes

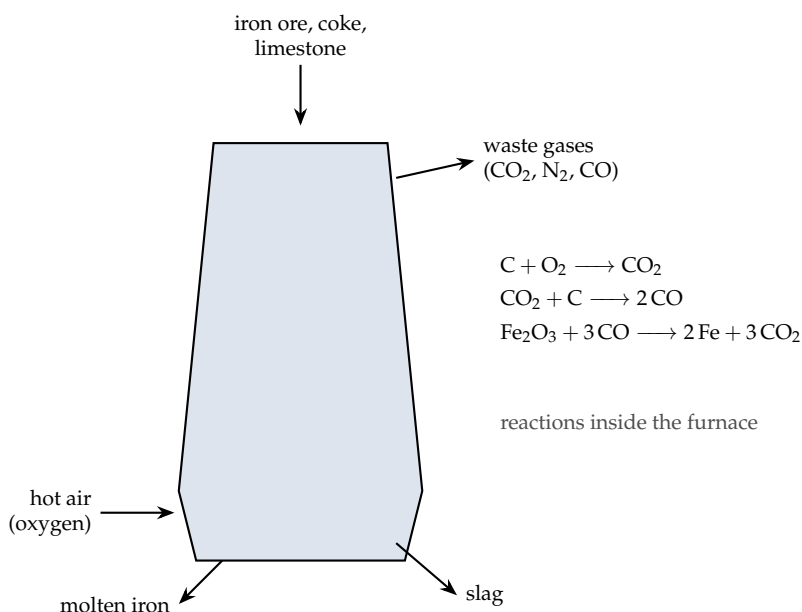
- Putting carbon or hydrogen in the wrong place: both are reference points, with carbon between aluminium and zinc and hydrogen between iron and copper.
- Predicting that a metal reacts with a compound of a more reactive metal: Cu does not react with ZnSO_4 .
- Describing the ionic equation with unbalanced charge: in $\text{Zn} + \text{Cu}^{2+} \longrightarrow \text{Zn}^{2+} + \text{Cu}$ the charges are +2 on both sides.

5 Extracting metals from their ores

Ores. A metal is rarely found as the pure element (an exception is gold, found uncombined; silver also occurs native). Most metals are found as compounds, mostly oxides, in rocks called **ores**. Extraction is the conversion of the metal ion back into the metal atom: this is **reduction**, because the metal gains electrons (or loses oxygen).

Which method? It depends on the position in the reactivity series. Metals below carbon, such as iron and zinc, can be extracted by heating the oxide with carbon (or carbon monoxide), which takes the oxygen away. Metals above carbon, such as aluminium, bind to oxygen too strongly for carbon to remove it, so they are extracted by electrolysis of the molten compound (see the electrolysis lesson), which is expensive because it uses a great deal of electricity.

Iron in the blast furnace. Iron ore is mostly iron(III) oxide, Fe_2O_3 . Coke (carbon), the ore and limestone are fed in at the top and hot air is blown in at the bottom.



Carbon burns to make carbon dioxide (and heat), which is reduced by more coke to carbon monoxide, and carbon monoxide then reduces the iron oxide. The iron oxide loses oxygen, so it is reduced, and the carbon monoxide is oxidised:



The limestone (CaCO_3) decomposes to CaO , which combines with silica impurities to give slag (CaSiO_3); the molten iron and the slag are tapped off at the bottom.

Worked example: how much iron?

$\text{Fe}_2\text{O}_3 + 3\text{CO} \longrightarrow 2\text{Fe} + 3\text{CO}_2$. What mass of iron can be obtained from 320 t of iron(III) oxide, assuming complete conversion? (A_r : Fe = 56, O = 16.)

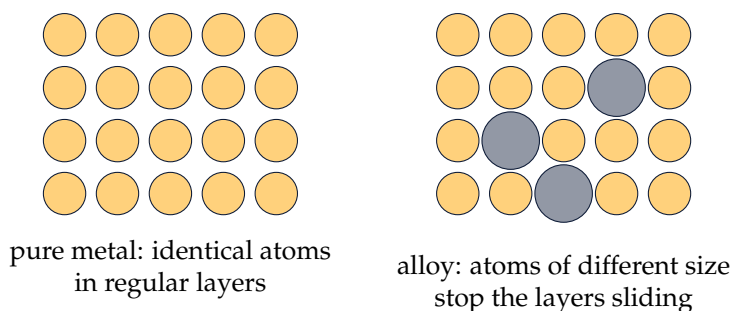
$M_r(\text{Fe}_2\text{O}_3) = 2 \times 56 + 3 \times 16 = 160$. So 320 t is $320/160 = 2$ units of 160. Each 160 of oxide gives $2 \times 56 = 112$ of iron, so the mass of iron is $2 \times 112 = 224$ t.

Common mistakes

- Saying that iron is extracted by electrolysis (that is for aluminium). Iron is below carbon and is reduced with carbon monoxide made from coke.
- Describing the conversion of the ore to metal as oxidation. The metal *gains* electrons or loses oxygen: reduction.
- Thinking that limestone is a reducing agent. It removes impurities (silica) as slag.

6 Uses of metals and alloys

Pure metals are often too soft. In a pure metal the atoms are all the same size and sit in regular layers that slide over each other easily. An **alloy** is a mixture of a metal with one or more other elements (metals or carbon). The different-sized atoms distort the layers, so they cannot slide, and the alloy is *harder and stronger* than the pure metal.



Uses. Each use comes from the properties:

metal	key properties	uses
aluminium	low density, resists corrosion (oxide layer), good conductor	aircraft, drinks cans, overhead power cables
iron / steel	strong, cheap, easy to shape; steel is an iron-carbon alloy; stainless steel contains chromium (and nickel) and does not rust	bridges, cars, cutlery
copper	very good electrical conductor, malleable, ductile, unreactive	electrical wiring, water pipes
silver	best electrical conductor, unreactive	jewellery, electrical contacts
gold	very unreactive, malleable, does not tarnish	jewellery, electronics
titanium	low density, very strong, resists corrosion	aircraft, replacement hip joints

Iron rusts. Iron rusts (reacts with oxygen and water) and must be protected by painting, oiling, galvanising (coating with zinc) or by alloying it into stainless steel.

Worked example: choosing a metal

Explain why aluminium is used for overhead power cables while copper is used for the wiring in buildings. Both conduct well. Aluminium is chosen for long overhead cables because of its low density (the cables are light enough to hang between pylons) and corrosion resistance. In buildings weight does not matter, and copper is a better conductor, is easier to bend and join, and does not corrode.

- Saying that an alloy is a compound. It is a mixture; there are no new chemical bonds, and its properties are those of a mixture of the atoms.
- Explaining hardness by “stronger bonds”. The reason is that layers cannot slide, because the atoms are of different sizes.
- Saying aluminium is unreactive. It is protected by an oxide layer (a thin, strongly adhering film) that stops further reaction.

Where are they? The transition metals sit in the central block of the Periodic Table, between Group 2 and Group 13 (groups 3–12): iron, copper, zinc, nickel, chromium, manganese, titanium, silver, gold and others.

- They form **ions with different charges** ("stable ions in several oxidation states"): iron forms Fe^{2+} and Fe^{3+} , copper forms Cu^+ and Cu^{2+} . This is why their names carry Roman numerals.
- Their compounds are often **coloured**: copper(II) compounds are blue, iron(II) compounds are pale green and iron(III) compounds are orange-brown, whereas Group 1 and Group 2 compounds are generally white or colourless.
- They (and their compounds) are often useful **catalysts**: iron in the Haber process for ammonia, and nickel when hydrogen is added to vegetable oils.

Worked example: transition metal or not?

A metal is hard and dense, melts at 1538 °C, forms a pale green chloride and an orange-brown chloride of different formulae, and acts as a catalyst. Is it a Group 1 metal or a transition metal?
The very high melting point, hardness and density, coloured compounds, two different ions (two different chlorides) and catalytic action are all transition metal features (this is iron: FeCl₂ and FeCl₃). A Group 1 metal is soft, melts below 200 °C, has one ion type (1+) and white compounds.

- Assuming all metals have the same pattern. Transition metals differ from Group 1: hard, dense, high melting points, less reactive.
- Giving the charge of an iron ion without checking the Roman numeral: iron(II) is Fe^{2+} , iron(III) is Fe^{3+} .
- Saying transition metals form white compounds. Typically they are coloured.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. Write the balanced equation for lithium with water.
2. Why is caesium more reactive than sodium?
3. Which halogen is a liquid at room temperature, and what colour is it?
4. Predict what happens when chlorine water is added to potassium iodide solution.
5. Why are noble gases unreactive?
6. Zinc is added to magnesium sulfate solution. Does it react? Explain.
7. Name the reducing agent that removes oxygen from iron oxide in the blast furnace.
8. Give two properties of transition metals that differ from those of Group 1 metals.

Answers to the self-check

1. $2\text{Li} + 2\text{H}_2\text{O} \longrightarrow 2\text{LiOH} + \text{H}_2$.
2. Its outer electron is in a shell further from the nucleus and more shielded, so it is lost more easily.
3. Bromine; red-brown.
4. Chlorine displaces iodine: the solution turns brown ($\text{Cl}_2 + 2\text{KI} \longrightarrow 2\text{KCl} + \text{I}_2$).
5. Their outer shells are full, so they do not need to gain, lose or share electrons.
6. No. Zinc is less reactive than magnesium, so it cannot displace it.
7. Carbon monoxide (made from coke).
8. Any two of: harder, denser, higher melting points, less reactive; coloured compounds; ions with different charges; used as catalysts.

Next step

Heurisco's *ESAT Chemistry: Groups and Metals Practice Set* has twenty original exam-style questions on this lesson, with data tables, graphs and apparatus, in order of difficulty, with a full worked solution for each.