Chemistry (0620)

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Standard Cheat Sheet

The Practice Book
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1 Acids, bases and salts

KEY FORMULAS

Neutralisation

$$H^+ + OH^- \rightarrow H_2O$$

Use whenever an acid reacts with an alkali. The hydrogen ion from the acid and the hydroxide ion from the alkali combine to make water, which is why every acid plus alkali reaction is called neutralisation.

A basic oxide neutralised by an acid

$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

Use as the template for any basic (metal) oxide reacting with an acid: the products are always a salt and water only, with no gas. The acid fixes the salt family, so sulfuric acid here gives the sulfate.

Acid plus metal gives a salt and hydrogen

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Use as the template for a reactive metal reacting with a dilute acid. The products are always a salt and hydrogen gas, and the hydrogen pops with a lighted splint. Only metals above hydrogen in reactivity react this way, so copper, silver and gold do not.

KEY CONCEPTS

- **Acids, bases and alkalis defined by their ions**: An *acid* produces hydrogen ions, H^+, when dissolved in water, and that one ion causes every acidic property. A *base* is a metal oxide or metal hydroxide that neutralises an acid. An *alkali* is a base that is soluble in water and gives hydroxide ions, OH^-. Every alkali is a base, but an insoluble base such as copper(II) oxide is not an alkali.
- **Basic and acidic oxides**: A *basic oxide* is a metal oxide that reacts with acids to give a salt and water, such as CaO and CuO. An *acidic oxide* is a non-metal oxide that reacts with bases or alkalis to give a salt and water, such as CO_2 and SO_2. The rule is that metal oxides are basic and non-metal oxides are acidic.
- **Indicator colours and the pH scale**: Litmus is red in acid and blue in alkali; thymolphthalein is colourless in acid and blue in alkali; methyl orange is red in acid and yellow in alkali. Universal indicator is matched to the pH scale, a number line from 0 to 14: below 7 is acidic, 7 is neutral and above 7 is alkaline.
- **Solubility rules for common salts**: All sodium, potassium and ammonium salts are soluble, and all nitrates are soluble. Most chlorides are soluble except silver chloride and lead(II) chloride. Most sulfates are soluble except barium sulfate, calcium sulfate and lead(II) sulfate. Most carbonates and most hydroxides are insoluble.
- **The three reactions of dilute acids**: A dilute acid reacts in three set patterns. Acid plus a reactive *metal* gives salt plus hydrogen. Acid plus a *base* (metal oxide or hydroxide) gives salt plus water. Acid plus a *carbonate* gives salt plus water plus carbon dioxide. Learn the patterns and the products follow automatically.
- **Bases, alkalis and the ammonia test**: A *base* is a metal oxide or hydroxide that neutralises an acid; the soluble ones are *alkalis* and give OH^- ions. Warming any ammonium salt with a base releases ammonia gas, which turns damp red litmus paper blue. This doubles as the standard test for an ammonium salt.
- **Making a soluble salt: excess-solid and titration**: For a soluble salt whose base is *insoluble* (a metal oxide, carbonate or reactive metal), add the solid in excess to the warm acid, filter off the unreacted excess, then evaporate to the point of crystallisation. For a soluble salt whose base is a *soluble alkali*, an excess cannot be filtered out, so find the exact neutralising volume by titration, then repeat with no indicator and crystallise.

EXAM TIPS

TIP Decide first whether the salt is soluble or insoluble, because that single choice picks the whole method. A *soluble* salt is made by reacting an acid with an excess insoluble base, carbonate or metal and then crystallising, or by titration when the base is a soluble alkali. An *insoluble* salt is made by precipitation, mixing two soluble solutions and filtering off the solid.

TIP The acid decides the salt's family and the other reactant supplies the metal. Hydrochloric acid gives a *chloride*, sulfuric acid gives a *sulfate* and nitric acid gives a *nitrate*. Name the salt by joining the metal to the acid's ending.

TIP An observation is something you can see, hear or feel, such as fizzing, a solid dissolving or a colour change. Naming a product, for example "carbon dioxide is given off", is a deduction and earns no observation mark. State the visible effervescence, not the gas.

2 Atoms, elements and compounds

KEY FORMULAS

Number of neutrons in an atom

$$\text{number of neutrons} = A - Z$$

Use to find the neutron count from the nucleon number A and the proton number Z. Protons and neutrons both sit in the nucleus, so subtracting the protons from the nucleon total leaves the neutrons.

Relative atomic mass from isotopic abundances

$$A_r = \frac{\sum(\text{isotopic mass} \times \%\ \text{abundance})}{100}$$

Use when an element exists as two or more isotopes with known percentage abundances. The result is a *weighted* mean, so it always lies closer to the more abundant isotope and need not be a whole number.

Deducing an ionic formula by charge balance

$$(\text{charge on cation}) \times a = (\text{charge on anion}) \times b$$

Use to find the whole-number ratio $a : b$ of ions in a compound, which is overall neutral. Combine the ions so the total positive charge cancels the total negative charge, then reduce to the lowest terms, as in $\mathrm{MgCl_2}$.

Two-isotope abundance equation

$$m_1 x + m_2(100 - x) = 100A_r$$

Use when an element has two isotopes of masses m_1 and m_2 and you must find the percentage x of the first from a known A_r. Rearrange to make x the subject; the larger percentage belongs to the isotope whose mass lies nearer the A_r.

KEY CONCEPTS

- **Covalent bonding**: A *covalent bond* is a shared pair of electrons between two atoms. It forms between non-metal atoms, where each shared pair lets both atoms count those electrons towards a full outer shell. A *simple molecular* substance is made of small separate molecules held to each other by weak intermolecular forces.
- **Element, compound and mixture**: An *element* is a substance made of only one type of atom and cannot be broken down into anything simpler by chemical means. A *compound* is two or more elements chemically combined in a fixed ratio, with new properties of its own. A *mixture* is two or more substances that are not chemically joined, present in any ratio, each keeping its own properties.
- **Giant covalent structures**: A *giant covalent* (macromolecular) structure is a continuous network of atoms joined throughout by strong covalent bonds, with no small separate molecules. Diamond and graphite are the required Core examples. Breaking the lattice means breaking many strong covalent bonds, so melting points are very high.
- **Ions and ionic bonding**: An *ion* is a charged particle formed when an atom loses or gains electrons. *Ionic bonding* is the strong electrostatic attraction between oppositely charged ions. It forms between a metal, which loses electrons to become a positive ion, and a non-metal, which gains them to become a negative ion, building a giant ionic lattice.
- **The three subatomic particles**: A proton has relative mass 1 and charge $+1$, a neutron relative mass 1 and no charge, and an electron negligible mass and charge -1. Protons and neutrons form the central nucleus while electrons occupy shells around it, so a neutral atom contains equal numbers of protons and electrons.

- **What isotopes are**: *Isotopes* are atoms of the same element with the same number of protons but different numbers of neutrons. They have identical electronic configurations, so their chemical properties are the same; they differ only in physical properties that depend on mass, such as density.
- **Conductivity of diamond and graphite**: In diamond each carbon atom bonds to four others, so every outer electron is held in a bond and diamond does not conduct electricity. In graphite each carbon bonds to only three others, leaving one delocalised electron per atom free to move along the layers, so graphite conducts.
- **Electronic configuration and position**: The *electronic configuration* lists the electrons in each shell from the inside out, filling 2 then 8 then 8 up to proton number 20, for example $2, 8, 1$ for sodium. The period number equals the number of occupied shells and the group number equals the number of outer-shell electrons.
- **Properties of ionic compounds**: An ionic compound has a high melting point because a large amount of energy is needed to overcome the strong electrostatic forces throughout the giant lattice. It conducts electricity when molten or dissolved, where the ions are free to move, but not when solid, where the ions are locked in fixed positions.
- **Properties of simple molecular substances**: Simple molecular substances have low melting and boiling points because melting or boiling overcomes only the weak intermolecular forces, not the strong covalent bonds inside the molecules. They do not conduct electricity, since the molecules are neutral with no free ions or electrons to carry charge.

EXAM TIPS

TIP To tell a compound from a mixture, ask how it is separated. A mixture is split by physical means such as filtration or distillation, while a compound can be broken up only by a chemical reaction or by electrolysis.

TIP When explaining a low boiling point, refer to the weak *intermolecular forces* being overcome, never the covalent bonds. Boiling pulls whole molecules apart, while the strong covalent bonds inside each molecule stay intact.

3 Chemical energetics

KEY FORMULAS

Energy transferred to water

$$E = V \times 4.2 \times \Delta T$$

Use to find the thermal energy in joules transferred to the water. V is the volume of water in cm^3, 4.2 is the energy in joules to raise 1 cm^3 of water by 1 °C, and ΔT is the temperature change in °C. The answer is in joules.

Temperature change

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

Use to find the temperature change of the surroundings. Subtract the initial reading from the final reading; a positive answer is a rise (exothermic) and a negative answer is a fall (endothermic). Keep both readings to the same number of decimal places.

Sharing a fixed energy through the water

$$\Delta T \propto \frac{1}{V}$$

Use when the same reaction releases the *same* amount of energy each time but the volume of water V is changed. The fixed energy shared through a larger volume gives a smaller temperature change, and through a smaller volume a larger one, so the temperature change is inversely proportional to the volume of water.

Temperature change from a known energy transfer

$$\Delta T = \frac{E}{V \times 4.2}$$

Use when the energy transferred E in joules is given and the temperature change is wanted. Rearranged from $E = V \times 4.2 \times \Delta T$, it divides the energy by the volume of water in cm^3 and by 4.2 to give the temperature change in °C.

KEY CONCEPTS

- **Common exothermic and endothermic changes**: Common *exothermic* changes are combustion of a fuel, neutralisation of an acid by an alkali, the reaction of a reactive metal with an acid, and respiration. Common *endothermic* changes are the thermal decomposition of a metal carbonate, photosynthesis, and the dissolving of ammonium nitrate in water.
- **Exothermic and endothermic reactions**: An *exothermic* reaction transfers thermal energy to the surroundings, so the temperature of the surroundings rises. An *endothermic* reaction takes thermal energy in from the surroundings, so the temperature of the surroundings falls. Energy cannot be seen, so the direction is always read from the surroundings: temperature up means exothermic, temperature down means endothermic.
- **Reaction pathway diagrams**: A *reaction pathway diagram* plots energy on the vertical axis against progress of reaction on the horizontal axis. If the products are drawn *below* the reactants the reaction released energy and is exothermic; if the products are drawn *above* the reactants the reaction took energy in and is endothermic. The vertical gap between the two levels is the energy transferred.
- **Comparing fuels or reactions fairly**: A fair comparison changes only the one thing being tested and keeps everything else the same. To compare fuels heating water, keep the same volume of water, the same starting temperature, the same distance from flame to container, and the same container; change only the fuel and measure one outcome, such as the time to reach a chosen temperature or the temperature rise in a fixed time.
- **Reducing heat loss**: Heat loss is reduced by wrapping *lagging* (insulation) around the container, fitting a *lid* to cut evaporation and rising warm air, and using an *insulated container* such as a polystyrene cup instead of a thin glass beaker. Each change keeps more energy in the water, so the measured temperature change moves closer to the true value.

- **The energy gap on a pathway diagram**: The vertical gap between the reactant level and the product level on a reaction pathway diagram is the amount of energy transferred: released in an exothermic reaction and taken in in an endothermic one. A *larger* gap means *more* energy is transferred, so of two exothermic reactions the one with the bigger downward step gives the bigger temperature rise.
- **Why heat loss makes a result too low**: A real experiment loses some thermal energy to the air, the bench, and the container, so less energy reaches the water. The measured temperature rise of an exothermic reaction is therefore *smaller* than the true value, and the measured fall of an endothermic reaction is smaller too. The chemistry is unchanged; only the measurement is reduced.

EXAM TIPS

TIP Classify a reaction by what the thermometer in the surroundings shows, not by the chemicals. A mixture that turns cold has not released *cold*; it has absorbed energy from its surroundings, so it is endothermic. Cold to the touch means endothermic every time, and warm means exothermic.

TIP A temperature change is always the final reading minus the initial reading. Adding the two readings together is the most common wrong answer in these questions, and quoting the final temperature on its own is the next. Write the subtraction first, before anything else.

TIP In the energy calculation put in the volume of *water* being warmed, never the mass of a solid that dissolved or the mass of fuel burned. The mass of dissolved solid is a distractor: it does not appear in $E = V \times 4.2 \times \Delta T$, and you never divide by it.

4 Chemical reactions

KEY FORMULAS

Mean rate of reaction

$$\text{mean rate} = \frac{\text{quantity of product formed (or reactant used up)}}{\text{time taken}}$$

Use to find the average rate over a measured interval, for example in cm^3 of gas per second or grams lost per second. Read the quantity from a gas syringe, a balance or a graph, then divide by the time taken.

Conservation of mass

$$\text{total mass of reactants} = \text{total mass of products}$$

Use because atoms are only rearranged in a reaction, never made or destroyed. An apparent gain in mass means a gas has been taken in from the air, and an apparent loss means a gas has escaped, for example when copper gains mass as it forms copper(II) oxide on heating.

Mean rate from a graph gradient

$$\text{rate} = \frac{\text{change in measured quantity}}{\text{change in time}}$$

Use to read a rate from a graph of product formed against time, where the steepness (gradient) of the curve gives the rate. The curve is steepest at the start, when most reactant is present, and becomes flat when the reaction has finished.

Relative rate of reaction

$$\text{relative rate} = \frac{1}{\text{time taken}}$$

Use to compare experiments timed by a fixed change, such as a cross disappearing behind a cloudy precipitate. It has no units, and a shorter time gives a larger relative rate, so the fastest run is always the one with the shortest time.

KEY CONCEPTS

- **Catalysts**: A *catalyst* is a substance that speeds up a reaction but is chemically unchanged at the end, so it can be recovered and used again. Its mass is the same before and after, and it changes only the *speed* of the reaction, never the amount of product formed.
- **Factors that affect the rate of reaction**: The rate of a reaction is increased by a higher *concentration* of a solution, a higher *temperature*, a larger *surface area* (smaller pieces or a powder) of a solid, and by adding a *catalyst*. A higher *pressure* also speeds up reactions between gases. Doing the opposite of any of these, such as diluting, cooling or using larger lumps, slows the reaction down.
- **Oxidation and reduction in terms of oxygen**: A *redox reaction* is one in which oxidation and reduction happen at the same time. At Core these are defined using oxygen: *oxidation* is the gain of oxygen and *reduction* is the loss of oxygen. They always occur together, because the oxygen lost by one substance is gained by another, as when a metal oxide is reduced while carbon is oxidised during the extraction of a metal.
- **Physical and chemical changes**: A *physical change* rearranges the same particles without making any new substance, for example melting, boiling, dissolving or crushing, and it is usually easy to reverse. A *chemical change* breaks and remakes bonds to form one or more *new substances* with different properties, signalled by a colour change, a gas, a precipitate or a permanent energy change, and it cannot be undone by simply cooling or evaporating.

- **Reversible reactions**: A *reversible reaction* can run both forwards and backwards, so the products can react together to re-form the reactants. It is shown by the double arrow $\rightleftharpoons$ in place of the single arrow $\rightarrow$, which marks a one-way reaction. Heating hydrated copper(II) sulfate to drive off its water, and then adding water again, is a reversible change.
- **Following the rate of a reaction**: A rate can be followed by measuring anything that changes as the reaction proceeds: the *volume of gas* collected in a gas syringe, the *loss of mass* on a balance as a gas escapes, or the *time* for a fixed change such as a marked cross disappearing behind a precipitate. Every method also needs a stop-watch, and the other conditions (especially temperature) must be kept constant for a fair comparison.
- **Hydrated and anhydrous salts (the test for water)**: A *hydrated* substance contains water of crystallisation, while an *anhydrous* substance has had that water removed. Hydrated copper(II) sulfate is *blue* and anhydrous copper(II) sulfate is *white*; hydrated cobalt(II) chloride is *pink* and anhydrous cobalt(II) chloride is *blue*. Adding water to white anhydrous copper(II) sulfate and seeing it turn blue is the chemical *test for water*.
- **Oxidation numbers and Roman numerals**: Some metals form ions of more than one charge, such as Fe^{2+} and Fe^{3+}, so the name carries a *Roman numeral in brackets* that gives the *oxidation number* of the metal. Iron(II) means an oxidation number of $+2$ and iron(III) means $+3$; copper(II) means $+2$. The numeral is not the number of atoms, and a higher value means the metal is more oxidised.
- **Reading a rate-time graph**: On a graph of product formed against time, the *steepness* of the curve is the rate and the *final height* of the plateau is the total amount of product. The curve is steepest at the start and flattens to horizontal when the reaction stops. A faster reaction is steeper and reaches its plateau sooner, but it can finish at the same height as a slower one when the amounts of reactant are the same.

EXAM TIPS

TIP If a solid appears to gain mass during a reaction, a gas from the air has joined it, as when an iron nail rusts. If a solid appears to lose mass, a gas has escaped into the air. In a sealed container the total mass never changes, because no gas can enter or leave.

TIP To show that a substance is *reduced*, point to oxygen disappearing from it between the left and right of the equation; to show it is *oxidised*, point to oxygen appearing on it. Quote the before-and-after formulae as your evidence. Keep the definitions the right way round: oxidation is *gain* of oxygen and reduction is *loss* of oxygen.

TIP Keep "how fast" and "how much" in separate boxes. Changing a rate condition such as concentration, surface area or a catalyst changes the *time* the reaction takes, not the *total* amount of product. The total is fixed by the amount of the reactant that is not in *excess*, so spotting the word "excess" tells you which reactant limits the product.

5 Chemistry of the environment

KEY FORMULAS

Complete combustion of carbon

$$C + O_2 \rightarrow CO_2$$

Use as the model for burning carbon or a carbon fuel in plenty of oxygen, where the only carbon product is carbon dioxide. This complete combustion is the main source of the carbon dioxide that acts as a greenhouse gas.

Combustion of sulfur impurities

$$S + O_2 \rightarrow SO_2$$

Use to show how the sulfur present in some fossil fuels burns to form sulfur dioxide. This gas dissolves in rain water to produce acid rain, which damages buildings, trees and aquatic life.

Incomplete combustion of carbon

$$2C + O_2 \rightarrow 2CO$$

Use to show what happens when a carbon fuel burns in too little oxygen. The product is carbon monoxide, a toxic gas that stops the blood carrying oxygen around the body.

KEY CONCEPTS

- **Chemical tests for the presence of water**: Two colour tests show that a liquid contains water. Anhydrous copper(II) sulfate turns from white to blue, and anhydrous cobalt(II) chloride turns from blue to pink. Both confirm only that water is *present*, not that it is pure, because any aqueous solution gives the same change.
- **Composition of clean, dry air**: Clean, dry air is about 78% nitrogen and 21% oxygen by volume, with about 1% argon and about 0.04% carbon dioxide. These are the gases of *normal* air. Any other gas, such as sulfur dioxide or carbon monoxide, is a *pollutant* and appears only in polluted air.
- **NPK fertilisers**: Crops remove nutrients from the soil, so farmers add fertilisers to replace them. An *NPK fertiliser* supplies nitrogen for healthy leaves, phosphorus for roots, and potassium for flowers and fruit. A single salt rarely holds all three elements, so an NPK product is usually a mixture of compounds.
- **The greenhouse effect and global warming**: *Greenhouse gases*, mainly carbon dioxide and methane, trap thermal energy that the warmed Earth gives out, keeping the lower atmosphere warmer. As fossil-fuel burning raises their concentration, more heat is trapped and the average global temperature rises. This is *global warming*, and it drives climate change.
- **Treating water to make it safe to drink**: The domestic supply is treated in three stages in a fixed order: *sedimentation* lets large insoluble particles settle out, *filtration* removes the finer suspended solids, and *chlorination* adds a small amount of chlorine to kill harmful microbes. The water is then safe to drink but still holds dissolved substances, so it is not pure.
- **Reducing air pollution**: Each pollutant has its own control method. A *catalytic converter* in a car exhaust changes harmful carbon monoxide and oxides of nitrogen into less harmful gases. *Flue gas desulfurisation* removes sulfur dioxide at a power station by neutralising the acidic gas with a base such as calcium oxide. *Particulate filters* trap the solid particles from diesel engines.
- **Sources and harms of the main air pollutants**: *Carbon monoxide* forms in incomplete combustion and is toxic because it stops the blood carrying oxygen. *Sulfur dioxide* forms when fuels containing sulfur burn and causes acid rain. *Oxides of nitrogen* form in hot engines and cause both acid rain and breathing problems. *Particulates*, the tiny solids from diesel engines, also harm the lungs.
- **Substances found in natural water**: Natural water is never pure. Helpful substances include dissolved oxygen, which aquatic life needs, and dissolved calcium compounds, which support shells and bones. Harmful substances include harmful microbes from sewage and nitrates and phosphates that run off from fertilisers and detergents.

EXAM TIPS

TIP Nitrogen makes up 78% of clean air yet is *not* a greenhouse gas, while carbon dioxide is only about 0.04% yet is a main cause of global warming. A gas warms the climate by trapping the heat the Earth gives out, so judge it by that property, never by how common it is.

TIP The copper(II) sulfate and cobalt(II) chloride colour changes prove only that water is present; a salt solution triggers them just as well. To show that water is *pure*, check for a sharp, fixed boiling point of 100 °C or melting point of 0 °C, because a dissolved impurity raises the boiling point above 100 °C and spreads it over a range.

TIP Sort each pollutant by the *kind* of harm it does. Carbon monoxide is toxic to the blood; sulfur dioxide and oxides of nitrogen cause acid rain; carbon dioxide and methane cause global warming. Matching a gas to the wrong harm, such as blaming sulfur dioxide for global warming, is the commonest slip in this topic.

TIP In an exam, work out the nutrient from the ions in the formula. The ammonium ion $\mathrm{NH_4^+}$ and the nitrate ion $\mathrm{NO_3^-}$ both supply nitrogen, the phosphate ion $\mathrm{PO_4^{3-}}$ supplies phosphorus, and the potassium ion $\mathrm{K^+}$ supplies potassium. A compound can therefore provide more than one of the N, P and K elements at once.

6 Electrochemistry

KEY FORMULAS

Decomposition of molten lead(II) bromide

$$PbBr_2 \rightarrow Pb + Br_2$$

Use to summarise the overall change when molten lead(II) bromide is electrolysed with inert electrodes. The lead collects as a silvery bead at the cathode and the bromine appears as a reddish-brown vapour at the anode.

Overall fuel-cell reaction

$$2H_2 + O_2 \rightarrow 2H_2O$$

Use to represent the overall reaction in a hydrogen-oxygen fuel cell, in which hydrogen and oxygen react to make water. Note that water is on the *right*, because the cell makes water; the reverse, with water on the left, is the electrolysis of water.

Electrolysis of water

$$2H_2O \rightarrow 2H_2 + O_2$$

Use to represent what the electrolysis of dilute sulfuric acid is effectively doing. Two volumes of hydrogen form for every one volume of oxygen, which is the source of the 2 to 1 ratio seen in the collecting tubes.

KEY CONCEPTS

- **Anode, cathode and electrolyte**: The liquid that is broken down is the *electrolyte*. The *cathode* is the electrode joined to the negative terminal and the *anode* the electrode joined to the positive terminal. The electrodes are usually *inert*, meaning they conduct the current but do not react; carbon (graphite) and platinum are the two standard inert materials.
- **How charge is carried**: Charge moves through the wires as a flow of *electrons*, exactly as in any metal circuit, but through the electrolyte as *moving ions*. Positive ions (cations) move to the cathode and negative ions (anions) move to the anode. Electrons never travel through the electrolyte and ions never travel along the wires.
- **Products of a molten binary compound**: A binary compound contains just two elements, a metal and a non-metal. When it is molten there is no water present, so the rule is simple: the *metal* forms at the cathode and the *non-metal* forms at the anode. Molten lead(II) bromide gives lead at the cathode and bromine at the anode.
- **The hydrogen-oxygen fuel cell**: A hydrogen-oxygen fuel cell combines hydrogen and oxygen to produce electrical energy directly, and the *only chemical product is water*. The hydrogen is the fuel that is used up and oxygen is the gas it reacts with. Unlike an ordinary battery, the reactant gases are supplied continuously from outside, so the cell keeps working as long as they are fed in.
- **What electrolysis is**: Electrolysis is the breakdown of an ionic compound, when *molten* or in *aqueous solution*, by the passage of an electric current. Each part is examined: it must be an *ionic* compound, it must be molten or dissolved so the ions are free to move, and the energy that drives it comes from the power supply. A covalent compound has no ions and cannot be electrolysed.
- **Electrolysis of concentrated sodium chloride**: Concentrated sodium chloride solution gives *hydrogen* at the cathode and *chlorine* at the anode, leaving sodium hydroxide in the solution. Hydrogen forms because sodium is more reactive than hydrogen, so the less reactive hydrogen ion is discharged. The chlorine can be tested with damp blue litmus paper, which turns red and is then bleached white.
- **Electrolysis of dilute sulfuric acid**: Dilute sulfuric acid is essentially acidified water, so its electrolysis effectively *splits water*. Hydrogen is released at the cathode and oxygen at the anode. Because water contains hydrogen and oxygen in a 2 to 1 ratio, the hydrogen is collected at twice the volume of the oxygen, and both gases are colourless.
- **Why a fuel cell pollutes less than a petrol engine**: Burning petrol in an engine releases carbon dioxide and other polluting gases, whereas a hydrogen-oxygen fuel cell produces only water at the

point of use. Because the cell converts the chemical energy of hydrogen straight into electrical energy and emits no carbon dioxide as it runs, it causes far less air pollution than a petrol engine.

- **Why a solid does not conduct but a melt does**: Conduction by an electrolyte needs charged particles that are free to *move*. In a solid ionic lattice the ions are locked in fixed positions, so no charge flows. Once the compound is melted, or dissolved in water, the lattice breaks down and the ions become free to move toward the electrodes. Melting does not create electrons or remove charges; it simply frees the ions that were already there.

EXAM TIPS

TIP Always decide the product from the *sign* of the electrode, not from the first letter of an element. The positive electrode (anode) gives the non-metal and the negative electrode (cathode) gives the metal or hydrogen. Guessing that electrode P gives lead because both start with the same letter is a common way to throw away a mark.

TIP In an aqueous solution the cathode product depends on reactivity. If the metal is *more reactive* than hydrogen, such as sodium or potassium, then hydrogen gas is released instead of the metal. If the metal is *less reactive* than hydrogen, such as copper or silver, then the metal itself is deposited.

7 Experimental techniques and chemical analysis

KEY CONCEPTS

- **Apparatus for measuring volume**: A *measuring cylinder* gives only an approximate volume. A *pipette* delivers one fixed accurate volume, such as $25.0\ \mathrm{cm}^3$. A *burette* delivers any chosen volume accurately, read to $0.1\ \mathrm{cm}^3$ through a tap, which is why it supplies the variable volume in a titration. A balance measures mass, not volume, so it can never answer a volume question.
- **Choosing a separation technique**: Match the method to what is mixed. *Filtration* separates an insoluble solid from a liquid, leaving the solid as residue and the liquid as filtrate. *Crystallisation* recovers a dissolved solid by evaporating to a saturated solution then cooling slowly. *Simple distillation* recovers a pure solvent, such as water, from a solution. *Fractional distillation* separates two miscible liquids with different boiling points, the lower-boiling liquid distilling over first.
- **Flame tests for metal cations**: A clean wire is dipped in the sample and held in a hot flame, giving a characteristic colour: lithium red, sodium yellow, potassium lilac, calcium orange-red, barium light green and copper(II) blue-green. Sodium yellow and calcium orange-red are easily confused, so the colours must be learned exactly.
- **Residue and filtrate**: When a mixture is filtered, the insoluble solid trapped on the paper is the *residue* and the liquid that passes through is the *filtrate*. The filtrate is not pure solvent, because anything already dissolved stays dissolved and passes straight through. A dissolved substance such as sugar cannot be filtered out, as its particles are far too small to be trapped by the paper.
- **Tests for common gases**: Hydrogen gives a squeaky pop with a lighted splint; oxygen relights a glowing splint; carbon dioxide turns limewater milky; ammonia turns damp red litmus paper blue; chlorine bleaches damp litmus paper. Each test follows the same pattern of add a test, observe a change, then state what it means.
- **The Rf value and how to measure it**: *Chromatography* separates dissolved substances such as the dyes in a food colouring. The *retardation factor* compares how far a substance travels with how far the solvent travels: $R_f = \dfrac{\text{distance moved by the substance}}{\text{distance moved by the solvent front}}$. Both distances are measured from the pencil baseline, the value has no units, and it always lies between 0 and 1. The same substance always gives the same R_f in the same solvent.
- **What a titration measures and the end-point**: A *titration* finds the volume of acid that exactly reacts with a fixed volume of alkali, so an unknown can be investigated. A pipette measures the alkali into a conical flask, a few drops of *indicator* are added, and acid is run in from a burette over a *white tile* until the colour just changes. That colour change is the *end-point*. *Methyl orange* is the usual indicator, turning from yellow in alkali to orange in acid.
- **Running a chromatogram correctly**: The baseline is drawn in *pencil*, because ink would dissolve and run with the solvent. The sample spots are placed on the baseline, which must sit *above* the solvent level so the spots are carried upward rather than washed off. The solvent is then allowed to rise up the paper. Two spots that rise to the same height share the same R_f and are likely the same substance, while several spots from one sample show it is a mixture.
- **Testing purity with melting and boiling points**: A *pure* substance melts and boils at one sharp, fixed temperature. An *impurity* makes the substance melt or boil over a *range* of temperatures instead of at a single point, and it lowers the melting point. So a solid that melts over a spread of temperatures is impure, while one that melts at a single fixed value is pure. Always state the reason, not only the verdict.
- **Tests for cations with sodium hydroxide**: Adding aqueous sodium hydroxide precipitates many metal hydroxides: copper(II) light blue, iron(II) green and iron(III) red-brown. Aluminium, zinc and calcium each give a *white* precipitate, so excess reagent is added to tell them apart: the aluminium and zinc precipitates redissolve in excess sodium hydroxide while calcium does not. Adding *excess ammonia* instead is the tiebreaker, as the zinc precipitate dissolves but the aluminium one does not.

EXAM TIPS

TIP Before testing for a halide or a sulfate, acidify the solution with dilute *nitric* acid to remove carbonate ions that would otherwise give a false positive. Hydrochloric acid must not be used for the halide test because it adds chloride ions, and sulfuric acid must not be used for the sulfate test because it adds sulfate ions.

TIP The *titre* is the volume of acid added, found as $\text{titre} = \text{final reading} - \text{first reading}$, never read straight off the burette. Both readings are taken to $0.1\ \text{cm}^3$. Repeat the titration and average only the *concordant* titres, those within about $0.10\ \text{cm}^3$ of each other, discarding the rough first run. *Universal indicator* is unsuitable because it passes through many colours and gives no single sharp end-point.

TIP Read a burette or measuring cylinder with the eye level with the *bottom of the meniscus*, the curved surface of the liquid. Looking from above or below introduces a *parallax* error and a wrong reading. A finer scale gives a smaller uncertainty, which is why a burette read to $0.1\ \text{cm}^3$ is more precise than a measuring cylinder.

8 Metals

KEY FORMULAS

Displacement of a less reactive metal

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

Use when a more reactive metal is added to a solution of a less reactive metal's salt. The more reactive metal takes the place of the less reactive one, which is deposited as a solid. Here the blue copper(II) sulfate fades and a pink-brown coating of copper forms on the zinc.

Reaction of a metal with dilute acid

$$\text{metal} + \text{dilute acid} \rightarrow \text{salt} + \text{hydrogen}$$

Use for any metal above hydrogen in the reactivity series, for example $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$. Bubbles of hydrogen are seen. Metals below hydrogen (copper, silver, gold) give no reaction with dilute acid.

Middle metal with steam

$$Zn + H_2O \rightarrow ZnO + H_2$$

Use for the middle metals (magnesium, zinc, iron), which react with *steam* rather than cold water to give a metal *oxide* plus hydrogen. These metals barely react with cold water.

Reaction of a metal with oxygen

$$2Mg + O_2 \rightarrow 2MgO$$

Use for a reactive metal burning in oxygen to form a *basic oxide*. Magnesium burns with a brilliant white flame to give white magnesium oxide. The more reactive the metal, the more vigorously it burns; unreactive metals barely tarnish.

Reactive metal with cold water

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

Use for the top three metals (potassium, sodium, calcium), which react with *cold water* to give a metal *hydroxide* plus hydrogen. The reaction is vigorous and the resulting solution is alkaline.

Reduction of iron(III) oxide in the blast furnace

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Use for the main reaction that extracts iron from hematite. Carbon monoxide is the reducing agent that removes oxygen from the ore. The ore is reduced (loses oxygen) and the carbon monoxide is oxidised (gains oxygen) to carbon dioxide.

KEY CONCEPTS

- **Conditions and product of rusting**: *Rusting* is the corrosion of iron and steel. It needs *both water and oxygen* present; salt speeds it up but is not essential. The product is *hydrated iron(III) oxide*, the orange-brown flaky solid. Remove either water or oxygen and rusting stops, which is the basis of every prevention method.
- **Extraction method from position relative to carbon**: A metal's extraction method follows its position relative to *carbon*. Metals below carbon (zinc, iron, copper) are extracted by *reduction with carbon*, the cheaper route, as iron is in the blast furnace. Metals above carbon (potassium to

aluminium) hold their oxygen too strongly and must be extracted by *electrolysis* of the molten compound. Very unreactive metals such as gold are found native.

- **Order of the reactivity series**: Metals are listed in order of how readily they react: *potassium, sodium, calcium, magnesium, aluminium*, [carbon], *zinc, iron*, [hydrogen], *copper, silver, gold*. Carbon and hydrogen are non-metal reference points. A higher metal reacts more vigorously, displaces a lower metal from a solution of its salt, and is harder to extract.
- **Physical and chemical properties of metals**: Most metals are *good conductors* of heat and electricity, *malleable* (can be hammered into shape), *ductile* (can be drawn into wires), shiny, dense and have high melting points. Chemically a metal reacts with dilute acid to give a salt and hydrogen, with water or steam to give hydrogen, and with oxygen to form a *basic oxide*. Non-metals are typically the reverse, so forming a basic oxide marks an element as a metal.
- **Uses of aluminium and copper**: *Aluminium* is used for aircraft and vehicle bodies, overhead power cables and food cans because it has a low density and resists corrosion through a tough surface oxide layer. *Copper* is used for electrical wiring and saucepan bases because it is an excellent electrical and thermal conductor, is ductile and does not react with water. A uses answer scores on linking the property to the job.
- **What an alloy is and why it is harder**: An *alloy* is a *mixture* of a metal with one or more other elements, usually other metals, so it is not a compound. It is harder and stronger than the pure metal because the added atoms are a *different size*, so they disrupt the regular layers of atoms and stop the layers sliding over one another. The metal atoms remain, so an alloy still conducts electricity.
- **Barrier methods of rust prevention**: Barrier methods keep water and oxygen *off* the iron: *painting* (railings, car bodies), *oiling or greasing* (moving parts), *coating with plastic* (wire fences) and *coating with another metal* such as zinc. Every barrier method shares one weakness: it only works while the coating is unbroken, so a scratch lets rusting begin.
- **Common alloys and their components**: *Brass* is copper with zinc (instruments, taps, fittings). *Bronze* is copper with tin (hard, corrosion-resistant). *Stainless steel* is iron with chromium and nickel (resists rusting; cutlery, sinks). *Mild steel* is iron with a little carbon (strong; construction). Each alloy is harder than the metal it is mainly based on.

EXAM TIPS

TIP When a question asks you to classify an unknown element, its oxide is the quickest test. A *metal* burns in oxygen to form a *basic oxide*, which reacts with acids; a *non-metal* forms an *acidic oxide*. Pair this with conducts plus malleable plus gives hydrogen with acid and the element is a metal every time.

TIP A uses question is marked on the *link*, not on the property name. Write the property, then because, then the consequence for the use, for example low density, so the aircraft is lighter, so it uses less fuel. A bare list of properties with no link to the job leaves marks unclaimed.

9 Organic chemistry

KEY FORMULAS

General formula of the alcohols

$$C_nH_{2n+1}OH$$

Use to write the formula of any alcohol from its carbon number n; for example ethanol has $n = 2$, giving C_2H_5OH. The *hydroxyl* group $-OH$ is the functional group, so every member shows the characteristic alcohol reactions.

General formula of the alkanes

$$C_nH_{2n+2}$$

Use to write the molecular formula of any alkane from its carbon number n; for example propane has $n = 3$, so its formula is C_3H_8. Alkanes are *saturated*, containing only single carbon-carbon bonds, which is why they hold the maximum possible number of hydrogen atoms and are generally unreactive.

General formula of the alkenes

$$C_nH_{2n}$$

Use to write the molecular formula of any alkene from its carbon number n; for example propene has $n = 3$, so its formula is C_3H_6. Alkenes are *unsaturated*, carrying one carbon-carbon double bond $C = C$ and therefore two fewer hydrogen atoms than the matching alkane.

General formula of the carboxylic acids

$$C_nH_{2n+1}COOH$$

Use to write the formula of any carboxylic acid; for example ethanoic acid takes $n = 1$ to give CH_3COOH. The *carboxyl* group $-COOH$ is the functional group, and it makes these compounds *weak acids* that still show the typical acid reactions.

Complete combustion of a hydrocarbon

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Use as the model for the *complete* combustion of any hydrocarbon in a plentiful supply of oxygen, which always gives carbon dioxide and water only. Balance the carbons first, then the hydrogens, and finish with the oxygens.

Substitution of an alkane in ultraviolet light

$$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$$

Use for the *substitution* reaction in which one hydrogen atom of an alkane is replaced by one chlorine atom, here forming chloromethane. The reaction needs *ultraviolet light* to start it and gives a hydrogen halide, HCl, as the second product; note the product is not hydrogen.

KEY CONCEPTS

- **Addition polymerisation of ethene**: A *polymer* is a very large molecule built when many small *monomers* join together. In *addition polymerisation* many alkene monomers join by opening their $C = C$ double bonds and form one long chain with no other product. The Core example is *poly(ethene)*, made from ethene; a polymer is named by writing *poly* in front of the monomer name.

- **Homologous series and functional groups**: A *hydrocarbon* is a compound that contains carbon and hydrogen only. A *functional group* is the atom or group of atoms that gives a molecule its characteristic reactions, such as the $-\mathrm{OH}$ group of the alcohols. A *homologous series* is a family of compounds that share the same functional group and the same general formula, so the members react alike while their physical properties change steadily as the chain grows. The functional groups to recognise are $\mathrm{C=C}$ (alkenes), $-\mathrm{OH}$ (alcohols), $-\mathrm{COOH}$ (carboxylic acids) and none (alkanes).
- **How an organic name is built**: An organic name is built from a *stem* that counts the carbon atoms in the chain and a *suffix* that names the functional group. The stems are *meth* (1 carbon), *eth* (2) and *prop* (3); the suffixes are *-ane* (alkane), *-ene* (alkene), *-ol* (alcohol) and *-oic acid* (carboxylic acid). So ethanol is *eth* (2 carbons) plus *-ol* (alcohol). Reading the suffix tells you the family and the functional group at once.
- **Petroleum and fractional distillation**: *Petroleum* (crude oil) is a mixture of hydrocarbons of many different chain lengths and is almost useless until separated. *Fractional distillation* sorts it into *fractions* of similar boiling point in a column that is hot at the bottom and cool at the top. Each fraction rises until it is cool enough to condense at its own level. Going down the column the chains are longer, so boiling point and viscosity rise and the fractions become harder to ignite.
- **Testing for unsaturation with bromine water**: An *unsaturated* compound contains at least one $\mathrm{C=C}$ double bond, while a *saturated* compound has only single carbon-carbon bonds. The test is to shake the compound with *bromine water*: an alkene adds bromine across its double bond and decolourises the orange solution to colourless, whereas an alkane has no double bond and leaves the colour unchanged.
- **Reactions of ethanoic acid**: Ethanoic acid $\mathrm{CH_3COOH}$ is a *weak acid* that still shows every typical acid reaction, and its salts are named *ethanoates*. With a reactive metal it gives a salt and hydrogen; with a base it gives a salt and water; with a carbonate it gives a salt, water and carbon dioxide. The salt is named as the metal followed by *ethanoate*, for example sodium ethanoate.
- **The two routes to ethanol**: Ethanol is made by two routes. *Fermentation* breaks down renewable glucose using yeast at about 30 °C in the absence of air; it is a slow process giving impure ethanol that must be distilled. The *catalytic addition* of steam to ethene at about 300 °C with an acid catalyst is fast and gives pure ethanol directly, but it uses ethene from non-renewable petroleum.

EXAM TIPS

TIP The stem counts *all* the carbon atoms in the molecule, including any carbon inside the functional group. Ethanoic acid $\mathrm{CH_3COOH}$ has two carbons, the one in the methyl group and the one in the $-\mathrm{COOH}$ group, so the stem is *eth* and the name is ethanoic acid, not methanoic acid. Forgetting the carbon in the functional group is a common naming error.

10 States of matter

KEY CONCEPTS

- **Changes of state**: A *change of state* is a reversible physical change in which no new substance forms. Melting (solid to liquid) and boiling (liquid to gas) take in energy; freezing (liquid to solid) and condensation (gas to liquid) give out energy. *Sublimation* is the direct change of a solid to a gas, shown by solid carbon dioxide and iodine.
- **Diffusion and the random motion of particles**: *Diffusion* is the spreading of particles from a region of higher concentration to a region of lower concentration, caused by their random motion. No stirring or draught is needed because the particles are always moving. Diffusion is fastest in gases, slower in liquids and negligible in solids, because the particles are progressively closer together and slower moving.
- **Properties of the three states**: Every bulk property follows from the particle picture. A solid keeps a fixed shape because its particles are locked in place; liquids and gases flow because their particles can move. Solids and liquids keep a fixed volume and resist compression because their particles already touch, whereas a gas fills its container, has a low density and is easily compressed because of the large spaces between its particles.
- **The three states in the kinetic particle model**: The *kinetic particle model* states that all matter is made of tiny particles in constant motion. In a *solid* the particles touch in a regular arrangement and only vibrate about fixed positions. In a *liquid* they touch in a random arrangement and slide past one another. In a *gas* they are far apart in a random arrangement and move quickly in all directions.
- **Evaporation versus boiling**: Both processes turn a liquid into a gas, but *evaporation* occurs only at the surface and at any temperature below the boiling point, whereas *boiling* occurs throughout the whole liquid and only at the boiling point. Because the most energetic particles escape during evaporation, the liquid left behind cools, which is why sweating cools the body.
- **Why diffusion is faster in a gas than in a liquid**: Diffusion happens in both gases and liquids but is much faster in a gas. In a gas the particles move quickly and have large spaces between them, so a particle travels a long way before colliding and the particles mix quickly. In a liquid the particles are closer together and move more slowly, colliding much sooner, so they mix far more slowly.
- **Why the temperature stays constant during a change of state**: Temperature measures the average kinetic energy of the particles. While a substance melts or boils, the energy supplied is used to overcome the forces of attraction between the particles rather than to make them move faster, so the temperature stays constant until the change is complete. On freezing or condensing the attractions re-form and release energy, again holding the temperature steady.

EXAM TIPS

TIP To compare the rate of diffusion in two states, compare how the particles move. In a gas the particles move quickly and have large spaces between them, so they spread out and mix rapidly; in a liquid they are closer together and move more slowly, so they mix far more slowly. Quote *both* speed and spacing for full marks.

TIP Place the given temperature on a number line marked with the melting point and the boiling point. Below the melting point the substance is solid, between the two points it is liquid, and above the boiling point it is gas. Watch negative values: a boiling point of -183 °C lies below room temperature, so the substance is already a gas at 25 °C.

TIP For a fixed mass of gas, raising the temperature at constant pressure increases the volume, and raising the pressure at constant temperature decreases the volume. A reliable check is that anything letting the particles spread out, whether more heat or less squeezing, increases the volume. Volume and pressure move in opposite directions; volume and temperature move together.

11 Stoichiometry

KEY FORMULAS

Amount of substance from mass

$$n = \frac{m}{M}$$

Use to convert between a mass m in grams and an amount n in moles, where M is the molar mass in g/mol. This is the first move of almost every reacting-mass calculation.

Number of particles from amount

$$N = n \times (6.02 \times 10^{23})$$

Use to find how many atoms, molecules or ions are present, multiplying the amount n in moles by the Avogadro constant. Read the substance carefully, because one mole of O_2 holds twice as many atoms as molecules.

Relative formula mass

$$M_r = \sum A_r$$

Use to find the mass of a molecule or formula unit by adding the A_r of every atom shown, multiplying each by its subscript. For calcium hydroxide $Ca(OH)_2$ this gives $40 + 2(16) + 2(1) = 74$.

Concentration of a solution

$$c = \frac{n}{V}$$

Use to link the concentration c in mol/dm³ to the amount of solute n in moles and the solution volume V in dm³. Convert any volume from cm³ to dm³ first, since 1 dm³ = 1000 cm³.

Percentage by mass of an element

$$\%\ \text{mass} = \frac{\text{total } A_r \text{ of the element}}{M_r} \times 100$$

Use to find what fraction of a compound's mass comes from one element. Multiply the element's A_r by how many of its atoms appear, divide by the M_r of the whole compound, then multiply by 100.

KEY CONCEPTS

- **Deducing an ionic formula from charges**: An ionic compound is overall neutral, so its formula is the smallest whole-number ratio of ions whose positive and negative charges cancel. Combine the ions until the total positive charge equals the total negative charge, then reduce to lowest terms. Aluminium oxide is Al_2O_3 because two Al^{3+} (total 6+) balance three O^{2-} (total 6−).
- **Relative atomic, molecular and formula mass**: The *relative atomic mass* A_r is the average mass of an element's atoms on a scale where one carbon-12 atom is exactly 12. The *relative molecular mass* M_r is the sum of the A_r values of every atom in a molecule; for an ionic compound the same sum is called the *relative formula mass*. All three are ratios, so they have no units.
- **The mole and the Avogadro constant**: The *mole* is the chemist's counting unit: one mole of any substance contains 6.02×10^{23} particles, a number called the *Avogadro constant*. The mole is the hub of stoichiometry because mass, number of particles and concentration are each linked to one another only through the amount in moles.

- **Counting atoms with brackets**: A subscript outside a bracket multiplies *every* atom inside it, not just the first. In ammonium sulfate $(NH_4)_2SO_4$ the outer 2 doubles the whole NH_4 group, giving 2 N and 8 H, then the SO_4 adds 1 S and 4 O. Count the bracketed group first, then the rest, then add.
- **Relative atomic mass as a weighted average**: Most elements exist as a mixture of *isotopes*, which are atoms of the same element with different masses. The relative atomic mass is the average of those isotope masses, weighted by how common each one is, so it lies closer to the more abundant isotope. This is why chlorine has $A_r = 35.5$ rather than a whole number.
- **The three-move reacting-mass route**: Every reacting-mass calculation follows one pattern: *get to moles* of what you are given with $n = m/M$, *cross the equation* by multiplying by the ratio of the balancing numbers, then *leave moles* by converting to what the question asks for. The only chemistry step is the middle one; the rest is unit conversion.

EXAM TIPS

TIP To balance a symbol equation, adjust only the *balancing numbers* written in front of each formula, never the subscripts inside a formula. The number of atoms of every element must then be equal on both sides. Changing a subscript changes the substance itself and scores no marks.

TIP When a polyatomic ion such as hydroxide OH^-, nitrate NO_3^- or sulfate SO_4^{2-} is needed more than once, wrap it in brackets so the subscript multiplies the whole group. Magnesium hydroxide is $Mg(OH)_2$, never $MgOH_2$, and iron(III) sulfate is $Fe_2(SO_4)_3$.

TIP Atoms are only rearranged in a reaction, so the total mass of products equals the total mass of reactants. Because a fixed reaction scales as a whole, halving the mass of a reactant halves the mass of every product. An apparent loss or gain of mass means a gas has escaped or been absorbed.

12 The Periodic Table

KEY FORMULAS

Group I metal reacting with water

$$2\mathrm{M} + 2\mathrm{H_2O} \rightarrow 2\mathrm{MOH} + \mathrm{H_2}$$

Use for any alkali metal M reacting with water. The products are a soluble metal hydroxide, which makes the solution alkaline, and hydrogen gas. The gas released is hydrogen, never oxygen.

Ion charge predicted from Group number

$$\mathrm{I} \rightarrow 1+, \quad \mathrm{II} \rightarrow 2+, \quad \mathrm{III} \rightarrow 3+, \quad \mathrm{VI} \rightarrow 2-, \quad \mathrm{VII} \rightarrow 1-$$

Use to predict the charge on a simple ion directly from its Group. A metal in Groups I to III loses electrons to give a positive charge equal to the Group number; a non-metal in Groups VI and VII gains electrons to give a negative charge equal to eight minus the Group number.

Building a formula by balancing charges

$$\mathrm{M^{2+}} + 2\mathrm{X^{-}} \rightarrow \mathrm{MX_2}$$

Use to build the formula of a compound from the ion charges so the positive and negative charges cancel. The size of one ion's charge becomes the subscript of the other, then reduce to the lowest whole-number ratio. A 2+ ion combined with 1- ions needs two of them, giving $\mathrm{MX_2}$.

Halogen displacement equation

$$\mathrm{Cl_2} + 2\mathrm{KBr} \rightarrow 2\mathrm{KCl} + \mathrm{Br_2}$$

Use when a more reactive halogen is added to a solution of a less reactive halide. Chlorine takes the place of bromine in potassium bromide, releasing bromine, which turns the solution orange. The same pattern applies to chlorine added to potassium iodide, releasing iodine.

KEY CONCEPTS

- **Group I: the alkali metals**: The Group I elements (lithium, sodium, potassium and below) are the *alkali metals*. Each atom has one outer-shell electron and loses it to form a 1+ ion, so the Group has a single fixed valency. They are soft enough to cut with a knife and have low densities, so the lighter members float on water.
- **Group VII: the halogens**: The Group VII elements (fluorine, chlorine, bromine, iodine) are the *halogens*: reactive non-metals that exist as *diatomic* molecules such as $\mathrm{Cl_2}$. Each atom has seven outer electrons and gains one more to form a 1- *halide* ion such as $\mathrm{Cl^-}$.
- **Metals on the left, non-metals on the right**: Metals occupy the large region on the left of the table and non-metals the smaller region on the upper right. Position predicts the sign of the ion: a metal has few outer electrons and loses them to form a positive ion, while a non-metal has a nearly full outer shell and gains electrons to form a negative ion.
- **Noble gases: a full outer shell**: The Group VIII (Group 0) elements (helium, neon, argon and below) are the *noble gases*: colourless, *monatomic* gases that are chemically unreactive. Their atoms already have a complete outer shell (two electrons for helium, eight for the others), so they have no tendency to gain, lose or share electrons.
- **Reading the Periodic Table: Groups and Periods**: The elements are arranged in order of increasing proton number. A *Group* is a vertical column, and for Groups I to VII the Group number equals the number of outer-shell electrons, which is why a Group shares similar chemical properties. A *Period* is a horizontal row, and the Period number equals the number of occupied electron shells.
- **Transition elements: the defining properties**: The *transition elements* form the central block of the Periodic Table between Group II and Group III. Compared with the Group I metals they are hard,

dense and have high melting points, and they show two further signatures: they form *coloured compounds*, and they act as *catalysts*, such as iron in the Haber process.

- **Reactivity order and displacement in Group VII**: Reactivity *decreases* down Group VII, so the order is chlorine, then bromine, then iodine. A more reactive halogen *displaces* a less reactive halide from its salt: chlorine displaces both bromide and iodide, bromine displaces only iodide, and iodine displaces neither.
- **Transition elements versus Group I metals**: A genuine difference between a transition metal and a Group I metal must use a transition speciality, not a property they share. Both conduct electricity, so conduction scores no marks on a difference question. The real differences are that transition metals form *coloured* compounds rather than white ones, and that they are harder, denser and higher-melting and can act as catalysts.

EXAM TIPS

TIP When an alkali metal reacts with water the gas released is *hydrogen*, not oxygen. The reaction is metal plus water giving metal hydroxide plus hydrogen, so the hydrogen is displaced from the water. The hydroxide solution formed is alkaline and turns universal indicator purple.

TIP Other atoms react in order to reach a full, stable outer shell like that of a noble gas; the noble gases already have one, so they have no reason to react. This is why a sodium atom loses one electron to copy neon and a chlorine atom gains one to copy argon.