Chemistry (0620)

IGCSE • Extended • CAIE

Standard Cheat Sheet

The Practice Book
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1 Acids, bases and salts

KEY FORMULAS

Neutralisation as an ionic equation

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Use for every acid plus alkali reaction. The hydrogen ion from the acid and the hydroxide ion from the alkali combine to form water, and the spectator ions are left over as the dissolved salt.

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.

A soluble salt from an insoluble base

$$ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$$

Use as the equation for the excess-solid route to a soluble salt. An excess of the insoluble base zinc oxide neutralises all the acid, and the unreacted excess is later filtered off before crystallising the zinc sulfate.

Dissociation of a strong acid

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Use to show a *strong* acid in water. A single forward arrow signals full dissociation, so essentially every molecule splits into ions and the solution holds the maximum concentration of H^+.

KEY CONCEPTS

- **Acids, bases and alkalis defined by their ions**: An *acid* produces hydrogen ions, H^+, when dissolved in water, and this single 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, releasing 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 to give a salt and water, such as CO_2 and SO_2. As a rule, metal oxides are basic and non-metal oxides are acidic.
- **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 pH scale and indicator colours**: The pH scale runs from below 0 to 14: below 7 is acidic, 7 is neutral, above 7 is alkaline, and pH falls as the hydrogen-ion concentration rises. 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.
- **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.
- **Acid strength is not concentration**: *Strength* is the degree to which an acid dissociates into ions, a fixed property of the acid: strong acids fully dissociate, weak acids only partially dissociate.

Concentration is how much acid is dissolved per unit volume and changes on dilution. A weak acid can be concentrated and a strong acid can be dilute.

- **Excess-solid and titration routes**: For a soluble salt whose base is *insoluble* (a metal oxide, carbonate or unreactive-enough metal), add the solid in excess, filter off the excess, then crystallise. 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.
- **Oxide character across a period**: Reading left to right across a period, the oxides change from *basic* in the reactive metals, through *amphoteric* at the metal and non-metal borderline, to *acidic* in the non-metals. Knowing only whether the element is a metal or a non-metal usually predicts the oxide's class.

EXAM TIPS

TIP Decide first whether the salt is soluble or insoluble. 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 combining the metal with the acid's ending.

TIP An observation must be 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 scores 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

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

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 lowest terms, as in Al_2O_3.

Nuclide notation

$${}^{A}_{Z}\text{X}$$

Use to show the make-up of a nuclide, with the nucleon number A at the top-left and the proton number Z at the bottom-left of the symbol. For example ${}^{23}_{11}\text{Na}$ has 11 protons, 11 electrons and $23 - 11 = 12$ neutrons.

Two-isotope abundance equation

$$m_1 x + m_2(1 - x) = A_r$$

Use when an element has two isotopes of masses m_1 and m_2 and you must find the fraction x of the first from a known A_r. Rearrange to make x the subject, then multiply by 100 to give a percentage abundance.

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, graphite and silicon(IV) oxide are the required 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.

- **Metallic bonding**: *Metallic bonding* is the strong electrostatic attraction between a lattice of positive metal ions and a sea of delocalised electrons that move freely throughout the structure. Those delocalised electrons explain why metals conduct electricity even when solid and why their melting points are high.
- **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 (about $\frac{1}{1840}$) 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, graphite and silicon(IV) oxide**: In diamond each carbon bonds to four others and in silicon(IV) oxide every outer electron is held in a bond, so neither conducts electricity. In graphite each carbon bonds to only three others, leaving one delocalised electron per carbon 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.
- **Malleability and conductivity of metals**: Metals are malleable and ductile because layers of positive ions can slide over one another while the sea of delocalised electrons keeps holding them together, so the metal changes shape without shattering. The same delocalised electrons are free to move and carry charge, which makes metals good conductors of electricity.
- **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.
- **Pure substances and fixed composition**: A *pure substance* is a single element or compound with no other substance mixed in; it melts and boils at a sharp, fixed temperature. Every sample of a given compound has the same composition because its elements are combined in a fixed ratio by mass set by the formula, so the proportions never vary.

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.

TIP Because relative atomic mass is weighted by abundance, the value always sits nearer the mass of the more common isotope. A result that lands exactly halfway between the two isotope masses means the two abundances are equal.

3 Chemical energetics

KEY FORMULAS

Enthalpy change from bond energies

$$\Delta H = \sum E(\text{bonds broken}) - \sum E(\text{bonds formed})$$

Use to calculate an enthalpy change from a table of bond energies. The reactant bonds broken take energy in and are added; the product bonds formed give energy out and are subtracted. The subtraction supplies the correct sign of ΔH automatically.

Sign convention for the enthalpy change

$$\Delta H < 0 \text{ (exothermic)}, \quad \Delta H > 0 \text{ (endothermic)}$$

Use to read the type of reaction directly from the sign of ΔH. A negative value means the chemicals lose stored energy and release it (exothermic); a positive value means the chemicals gain stored energy from the surroundings (endothermic).

Counting bonds from the equation

$$n(\text{bond}) = (\text{coefficient}) \times (\text{number of that bond per formula})$$

Use to find how many of each bond break or form before applying the bond-energy equation. A coefficient multiplies every bond in that formula, so $2H_2O$ contains four O-H bonds, and a double or triple bond counts as one bond with one large bond energy.

Reverse reaction flips the sign

$$\Delta H_{\text{reverse}} = -\Delta H_{\text{forward}}$$

Use when a reaction is reversed, such as comparing respiration with photosynthesis. The reverse reaction transfers the same quantity of energy in the opposite direction, so the magnitude of ΔH is unchanged and only its sign reverses.

KEY CONCEPTS

- **Bond breaking and bond making**: *Bond breaking* requires energy to pull bonded atoms apart, so it is always endothermic. *Bond making* releases energy as atoms join, so it is always exothermic. Whether the whole reaction is exothermic or endothermic depends on which of these two energy transfers is larger.
- **Bond energy**: A *bond energy* is the energy needed to break one mole of a particular bond, which is exactly equal to the energy released when one mole of that same bond forms. One value therefore describes both breaking and making the bond, which is why a single table of bond energies can predict an enthalpy change.
- **Common exothermic and endothermic changes**: Common *exothermic* changes are the combustion of fuels, the neutralisation of an acid by an alkali, and the reaction of a reactive metal with an acid. Common *endothermic* changes are the thermal decomposition of a carbonate and the dissolving of ammonium nitrate or ammonium chloride in water.
- **Enthalpy change**: The *enthalpy change* ΔH is the energy transferred to or from the surroundings per mole of reaction, measured in kJ/mol. The symbol Δ means *change in* and H is the enthalpy, the energy stored in the chemicals. An exothermic reaction releases stored energy and a thermal decomposition is a common endothermic process.
- **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. The energy change is always judged from the surroundings, most usefully the water or solution holding the thermometer.

- **Activation energy**: The *activation energy* E_a is the minimum energy that colliding particles must have before they can react. It is the height of the barrier between reactants and products, so a reaction with a high activation energy is slow at room temperature because only a small fraction of collisions are energetic enough to succeed.
- **Reaction pathway diagrams**: A *reaction pathway diagram* plots energy against progress of reaction, with reactants on the left and products on the right. For an exothermic reaction the products are drawn *below* the reactants; for an endothermic reaction they are drawn *above*. The peak above the reactants is the activation energy, and the vertical gap between reactants and products is ΔH.

EXAM TIPS

TIP Energy cannot be seen, so classify a reaction by the surroundings, not the chemicals. If the surroundings get hotter the reaction is exothermic; if they get colder it is endothermic. A mixture that goes cold has not released *cold*; it has absorbed energy from its surroundings.

TIP When a mixture cools as it reacts, energy has flowed from the surroundings into the reaction, so the change is endothermic and ΔH is positive. Nothing releases *cold*; the falling temperature is the surroundings losing energy to the reacting chemicals.

TIP Read the balanced equation and the structure of every molecule before summing energies. Multiply each bond energy by how many of that bond are present, treating a double bond such as O=O or a triple bond such as N≡N as a *single* bond with its own value, never as two or three separate bonds.

4 Chemical reactions

KEY FORMULAS

Average rate of reaction

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

Use to find the mean 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 elapsed.

Oxidation number sum rule

$$\sum(\text{oxidation numbers}) = \text{charge on the species}$$

Use to find an unknown oxidation number. The oxidation numbers of all atoms add up to zero in a neutral compound, or to the overall charge in an ion. Treat oxygen as -2 and hydrogen as $+1$ in their usual compounds, and an uncombined element as 0.

Mean rate as a graph gradient

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

Use to read a rate from a graph of product formed against time, where the gradient of the curve gives the rate at any stage. The curve is steepest at the start, when reactant is most plentiful, and flattens to zero when the reaction stops.

KEY CONCEPTS

- **Collision theory**: *Collision theory* states that particles can only react when they collide, and the collision must have energy at least equal to the *activation energy* and the correct orientation. Most collisions are too gentle or badly aligned, so only a small fraction are successful; the rate of reaction is set by the number of successful collisions per second.
- **Conservation of mass**: In any chemical reaction atoms are only rearranged, never created or destroyed, so the total mass of the products equals the total mass of the reactants. An apparent gain in mass means a gas has been taken in from the surroundings, and an apparent loss means a gas has escaped into them.
- **Dynamic equilibrium**: A *reversible reaction*, shown by the sign $\rightleftharpoons$, reaches *dynamic equilibrium* in a closed system when the forward and reverse reactions occur at equal rates. At equilibrium the concentrations of reactants and products stay constant, though usually not equal, and both reactions are still taking place.
- **Oxidation and reduction (OIL RIG)**: A *redox reaction* is one in which oxidation and reduction happen together. *Oxidation* is the gain of oxygen, the loss of electrons, or an increase in oxidation number; *reduction* is the loss of oxygen, the gain of electrons, or a decrease in oxidation number. The mnemonic *OIL RIG* records the electron view: Oxidation Is Loss, Reduction Is Gain.
- **Physical and chemical changes**: A *physical change* rearranges the same particles without making any new substance, for example melting, boiling or dissolving, and it is usually easily reversed with no change in mass. 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 reversed by simple cooling or evaporation.
- **Factors that change the rate of reaction**: Five factors increase the rate of a reaction. Raising the *concentration* of a solution, the *pressure* of a gas, or the *surface area* of a solid each makes collisions more frequent. Raising the *temperature* makes collisions both more frequent and more energetic, which is why its effect is so large. Adding a *catalyst* lowers the activation energy, so a greater proportion of collisions succeed.
- **Le Chatelier's principle**: When a condition is changed, the position of equilibrium shifts so as to oppose that change. Adding a substance shifts the equilibrium away from it, while removing one

shifts it towards that side. Raising the pressure shifts a gaseous equilibrium towards the side with fewer moles of gas. Raising the temperature shifts it in the *endothermic* direction. A catalyst shifts the position not at all; it only makes equilibrium arrive sooner.

- **Oxidising agents and reducing agents**: An *oxidising agent* oxidises another substance and is itself reduced, taking electrons or giving oxygen. A *reducing agent* reduces another substance and is itself oxidised, giving electrons or taking oxygen. Each agent is named for its effect on its partner, so an oxidising agent is always the species that is itself reduced.

EXAM TIPS

TIP When explaining any rate change, first decide which lever has moved: the *frequency* of collisions (how often particles meet) or the *proportion* of collisions that are successful (how many reach the activation energy). Concentration, pressure and surface area change only the frequency; temperature changes both; a catalyst changes only the proportion.

TIP If a solid appears to gain mass during a reaction, a gas has been absorbed into it; a rusting nail gains mass as oxygen from the air joins the iron. If a solid appears to lose mass, a gas has escaped; a burning sparkler loses mass as gaseous products leave. In a sealed container the total mass never changes, because no gas can enter or leave.

TIP When predicting a pressure shift, count moles of *gas only* and ignore any solids and liquids. If both sides have equal moles of gas, a change in pressure has no effect on the position. Tie a temperature shift to whichever direction is *endothermic* rather than to "forwards" by habit, because that depends on the sign of the enthalpy change.

5 Chemistry of the environment

KEY FORMULAS

Complete combustion of carbon

$$C + O_2 \rightarrow CO_2$$

Use as the model for burning a fuel in plenty of oxygen, where the only carbon product is carbon dioxide. A sulfur impurity burns in the same way, $S + O_2 \rightarrow SO_2$, releasing the sulfur dioxide that goes on to cause acid rain.

Percentage by mass of an element

$$\%\,\text{element} = \frac{\text{mass of the element in one formula unit}}{M_r} \times 100$$

Use to compare fertilisers by the share of a nutrient they carry, most often the percentage of nitrogen. Count *every* atom of the element in the formula, so NH_4NO_3 contributes two nitrogen atoms, a mass of 28, not 14.

Making ammonium nitrate

$$NH_3 + HNO_3 \rightarrow NH_4NO_3$$

Use to show how a nitrogen fertiliser is made by neutralising nitric acid with ammonia. The ammonium ion NH_4^+ and the nitrate ion NO_3^- each carry a single charge, so they combine in a one to one ratio with no spare atoms.

Reaction inside a catalytic converter

$$2CO + 2NO \rightarrow 2CO_2 + N_2$$

Use to show how a catalytic converter removes two pollutants at once: carbon monoxide is oxidised to carbon dioxide and nitrogen monoxide is reduced to harmless nitrogen. It improves local air quality but still releases carbon dioxide, which is itself a greenhouse gas.

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, 21% oxygen and 1% argon by volume, with about 0.04% carbon dioxide and traces of other gases making up the rest. Nitrogen and oxygen together account for roughly 99%, so the remaining gases, though tiny in amount, include the carbon dioxide that drives climate change.
- **NPK fertilisers**: Fertilisers replace the elements that crops take from the soil. An *NPK fertiliser* supplies nitrogen for leafy growth, phosphorus for roots, and potassium for flowers and fruit. A single salt rarely holds all three, so an NPK product is usually a mixture, for example ammonium phosphate combined with potassium chloride.
- **The greenhouse effect and global warming**: The Earth's surface radiates thermal energy, *greenhouse gases* such as carbon dioxide and methane absorb it and re-emit some back towards the surface, and this keeps the lower atmosphere warmer than it would otherwise be. Rising concentrations of these gases trap more of this energy, raising the average global temperature, which is *global warming*.
- **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, controlled amount of chlorine to kill harmful microbes. The result is safe to drink but still holds dissolved salts, so it is not pure.

- **Eutrophication**: When nitrates and phosphates from fertiliser run-off or sewage over-enrich a river or lake, algae grow rapidly. As the algae die they are decomposed by microbes, the microbes use up the dissolved oxygen, and fish and other aquatic life then suffocate. This whole sequence is called *eutrophication*.
- **Sources and harms of the main air pollutants**: *Carbon monoxide* comes from incomplete combustion and is toxic because it binds to haemoglobin in place of oxygen. *Sulfur dioxide* comes from burning sulfur-containing fuels and dissolves in rain to cause acid rain. *Oxides of nitrogen* form in hot engines and cause both acid rain and smog. *Carbon dioxide* comes from complete combustion and drives global warming.

EXAM TIPS

TIP Nitrogen makes up 78% of air yet is *not* a greenhouse gas, while carbon dioxide is only about 0.04% yet is one of the main causes of global warming. A gas warms the climate by absorbing the thermal energy the Earth radiates, 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 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 and lowers the melting point.

TIP Sort each pollutant by the *kind* of harm it does. Carbon monoxide is toxic to the blood but causes neither acid rain nor global warming; sulfur dioxide and oxides of nitrogen cause acid rain; carbon dioxide and methane cause global warming. Matching a gas to the wrong harm is the commonest exam slip in this topic.

6 Electrochemistry

KEY FORMULAS

Anode half-equation for a halide ion

$$2Br^- \rightarrow Br_2 + 2e^-$$

Use at the anode, where a halide ion loses an electron and is oxidised to a halogen *molecule*. Two ions are always needed because the product is diatomic, for example Br_2 or Cl_2.

Cathode half-equation for a metal ion

$$Pb^{2+} + 2e^- \rightarrow Pb$$

Use at the cathode, where a metal cation gains electrons and is reduced to the neutral metal. The number of electrons equals the charge on the ion, so an Al^{3+} ion needs three electrons while a Pb^{2+} ion needs two.

Overall fuel-cell reaction

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

Use to summarise the overall change in a hydrogen-oxygen fuel cell, in which hydrogen and oxygen react to form water. It is exactly the reverse of the electrolysis of water, $2H_2O \rightarrow 2H_2 + O_2$.

Discharge of hydrogen ions

$$2H^+ + 2e^- \rightarrow H_2$$

Use at the cathode whenever hydrogen is the species discharged, for example in dilute acids or in solutions of a reactive-metal salt. Two ions gain two electrons to form one molecule of hydrogen gas.

Discharge of hydroxide ions

$$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$$

Use at the anode when oxygen is released, as in dilute salt solutions and dilute acids. Four hydroxide ions are oxidised to give water, oxygen and four electrons; this is the source of the oxygen seen bubbling at the anode.

KEY CONCEPTS

- **Electrolysis and the electrolytic cell**: Electrolysis is the breakdown of an ionic compound, when molten or in aqueous solution, by the passage of an electric current. The liquid decomposed is the *electrolyte*, and the current enters and leaves it through two *electrodes*. The *cathode* is joined to the negative terminal and the *anode* to the positive terminal. An *inert* electrode, such as carbon (graphite) or platinum, carries the current without itself reacting.
- **Oxidation at the anode, reduction at the cathode**: At the cathode, cations gain electrons, so reduction always occurs there; at the anode, anions lose electrons, so oxidation always occurs there. A useful memory aid is *red cat* (reduction at the cathode) and *an ox* (oxidation at the anode). Charge travels through the metal wires as moving electrons and through the electrolyte as moving ions.
- **Products from a molten binary compound**: A solid ionic compound does not conduct, because its ions are fixed in the lattice. Once molten the ions are free to move, so electrolysis can occur: the metal is discharged at the cathode and the non-metal at the anode. For molten lead(II) bromide this 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, with water as the only chemical product. Hydrogen is the *fuel* that is oxidised and oxygen is the *oxidising agent*. Unlike an ordinary battery, the reactant gases are supplied continuously from outside, so the cell runs as long as they are fed in.

- **Selective discharge at the anode**: At an inert anode in aqueous solution the choice lies between a halide ion and OH^- from water. A halide is discharged when its solution is *concentrated*, giving the halogen; in a dilute solution, or when no halide is present, hydroxide is discharged and oxygen is released. A stable oxoanion such as sulfate is never discharged.
- **Selective discharge at the cathode**: In an aqueous solution both a metal cation and H^+ from water reach the cathode, but only one is discharged. The *less reactive* species is discharged: a metal below hydrogen in the reactivity series, such as copper or silver, is deposited, while for a more reactive metal such as sodium the H^+ is discharged and hydrogen gas forms instead.
- **Why a fuel cell beats a petrol engine**: A fuel cell converts the chemical energy of hydrogen straight into electrical energy, avoiding the wasteful heat-engine step of a petrol engine, so a greater fraction of the energy becomes useful. It can be viewed as the electrolysis of water run in reverse, recombining hydrogen and oxygen to release electricity rather than using electricity to split water.

EXAM TIPS

TIP A correct half-equation balances both the atoms and the total charge. Add electrons to whichever side makes the charges equal: electrons appear on the right for an oxidation (loss) and on the left for a reduction (gain). Then check that the number of electrons matches the charge that has changed.

TIP A common error is to write $Cl^- \rightarrow Cl + e^-$. The product is the diatomic molecule Cl_2, so two ions must react: $2Cl^- \rightarrow Cl_2 + 2e^-$. The same applies to bromine and iodine, and to the oxygen and hydrogen released at the electrodes.

7 Experimental techniques and chemical analysis

KEY FORMULAS

Moles from concentration and volume

$$n = c \times V$$

Use to find the amount in moles n from a concentration c in $\mathrm{mol/dm^3}$ and a volume V that has first been converted to $\mathrm{dm^3}$. This is the opening step of every titration calculation.

Retardation factor

$$R_f = \frac{\text{distance moved by the substance}}{\text{distance moved by the solvent front}}$$

Use to identify a component on a chromatogram. Measure the substance distance from the pencil baseline to the centre of its spot, and the solvent distance to the solvent front. The value has no units and always lies between 0 and 1.

Concentration from moles and volume

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

Use to find a concentration c in $\mathrm{mol/dm^3}$ once the amount n in moles is known, with the volume V in $\mathrm{dm^3}$. This is the closing step of a titration, applied after the reacting ratio has converted the moles of one substance into moles of the other.

Concentration in grams per cubic decimetre

$$c\ (\mathrm{g/dm^3}) = c\ (\mathrm{mol/dm^3}) \times M_r$$

Use to convert a concentration from $\mathrm{mol/dm^3}$ to $\mathrm{g/dm^3}$ by multiplying by the relative formula mass M_r. The reverse conversion divides the mass concentration by M_r.

KEY CONCEPTS

- **Apparatus for measuring volume**: A *burette* delivers any chosen volume of liquid accurately, read to $0.1\ \mathrm{cm^3}$, and provides the variable volume in a titration. A *volumetric pipette* delivers one fixed accurate volume such as $25.0\ \mathrm{cm^3}$. A *measuring cylinder* is far less precise and is used only for approximate volumes. A balance measures mass, not volume, so it can never answer a volume question.
- **Choosing a separation technique**: *Filtration* separates an insoluble solid from a liquid, leaving the solid as the residue and the liquid as the filtrate. *Crystallisation* recovers a dissolved solid by evaporating to a saturated solution then cooling slowly. *Simple distillation* recovers a pure solvent from a solution, and *fractional distillation* separates two miscible liquids with different boiling points.
- **Flame tests for metal cations**: A clean wire dipped in the sample and held in a hot flame gives a characteristic colour: lithium red, sodium yellow, potassium lilac, calcium orange-red and copper(II) blue-green. Sodium yellow and calcium orange-red are deliberately set as look-alikes, as are potassium lilac and lithium red, so the five colours must be learned exactly.
- **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 blue; chlorine bleaches damp litmus paper; and sulfur dioxide turns acidified potassium manganate(VII) from purple to colourless.
- **What a titration measures**: A *titration* finds the volume of one solution that exactly reacts with a fixed volume of another, so an unknown concentration can be calculated. The *end-point* is where the

indicator changes colour, showing the acid and alkali have exactly reacted. Only *concordant* titres, agreeing to within about $0.10\ \mathrm{cm}^3$, are averaged; the rough first run is discarded.

- **Simple versus fractional distillation**: *Simple distillation* separates a solvent from a solution, for example pure water from sea water, by boiling off and condensing the solvent while the dissolved solute stays behind. *Fractional distillation* separates two miscible liquids whose boiling points differ; a *fractionating column* lets vapour repeatedly condense and re-evaporate, enriching it in the lower-boiling liquid so that liquid distils over first.
- **Stationary and mobile phases**: In paper chromatography the *stationary phase* is the chromatography paper and the *mobile phase* is the solvent that rises through it. The baseline is drawn in pencil because ink would dissolve and run with the solvent, and the sample spot must start above the solvent level so it is carried up rather than washed off. A *locating agent* such as ninhydrin reacts with colourless spots to make them visible.
- **Tests for anions**: To test for a *carbonate*, add dilute acid: effervescence gives a gas that turns limewater milky. For a *halide*, acidify with dilute nitric acid then add silver nitrate, giving white (chloride), cream (bromide) or yellow (iodide). For a *sulfate*, acidify with dilute nitric acid then add barium nitrate to give a white precipitate. For a *nitrate*, warm with sodium hydroxide and aluminium foil to release ammonia.
- **Tests for cations with sodium hydroxide**: Adding aqueous sodium hydroxide precipitates many metal hydroxides: copper(II) is light blue, iron(II) green and iron(III) red-brown, all insoluble in excess. A white precipitate that *redissolves* in excess sodium hydroxide is amphoteric aluminium hydroxide, while a white precipitate that stays is calcium hydroxide. The *ammonium* ion gives no precipitate but releases ammonia on warming.

EXAM TIPS

TIP Concentration is measured in $\mathrm{mol/dm}^3$, so each volume must be changed from cm^3 to dm^3 by dividing by 1000 before it is used in $n = c \times V$. Substituting a volume in cm^3 gives an answer one thousand times too large.

TIP Dilute *nitric* acid is added first to remove carbonate and sulfite ions that would otherwise give their own precipitate and 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 In a "which apparatus" question, first confirm the instrument measures the quantity asked for, then judge its precision. A balance, however precise, measures *mass* and can never answer a volume question. A pipette gives one fixed volume, while a burette gives an accurate variable volume read to $0.1\ \mathrm{cm}^3$.

8 Metals

KEY FORMULAS

Displacement of a less reactive metal

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

Use when a more reactive metal is placed in a solution of a less reactive metal's salt. The reactive metal forms ions and the less reactive metal is deposited. The anion, here SO_4^{2-}, is a spectator and is left out of the ionic equation.

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 $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$. The metal atom loses electrons to form a positive ion. Metals below hydrogen (copper, silver, gold) give no reaction.

Cathode reaction in aluminium extraction

$$Al^{3+} + 3e^- \rightarrow Al$$

Use for the negative electrode in the electrolysis of molten aluminium oxide. Aluminium ions gain three electrons each (reduction) and form molten aluminium that collects at the base of the cell.

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.

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 by carbon monoxide

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

Use for the main ore reduction in the blast furnace, where carbon monoxide is the reducing agent that removes oxygen from hematite. The ore is reduced (loses oxygen) and the carbon monoxide is oxidised (gains oxygen).

KEY CONCEPTS

- **Conditions and product of rusting**: *Rusting* is the corrosion of iron and steel. It needs *both water and oxygen* present; salt and acidity speed it up but are not essential. The product is hydrated iron(III) oxide, $Fe_2O_3 \cdot xH_2O$, the orange-brown flaky solid. Remove either water or oxygen and rusting stops.
- **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. 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.

- **Metallic bonding and the properties it explains**: A metal is a giant lattice of positive ions in a *sea of delocalised electrons* that move freely through the whole structure. The mobile electrons make a metal a good conductor of electricity and heat. The layers of ions slide without shattering, so a metal is *malleable* and *ductile*, and strong metallic bonding gives high melting and boiling points.
- **Order of the reactivity series**: Metals are listed in order of how readily their atoms lose electrons to form positive ions: *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 and is harder to extract.
- **Uses of aluminium and copper**: *Aluminium* is chosen for aircraft bodies, overhead power cables and food cans because it has a low density and resists corrosion through a self-sealing oxide layer. *Copper* is chosen for electrical wiring and saucepan bases because it is an excellent electrical and thermal conductor, is ductile, and does not react with water.
- **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. It is harder and stronger than the pure metal because the added atoms are a *different size*, so they disrupt the regular layers of ions and stop the layers sliding over one another. The metallic bonding remains, so an alloy still conducts.
- **Barrier methods and sacrificial protection**: Rusting is prevented in two ways. *Barrier methods* (painting, oiling, plastic coating, plating) keep water and oxygen off the iron but fail once scratched. *Sacrificial protection* attaches a *more reactive* metal such as zinc or magnesium, which corrodes in preference to the iron, so the iron is protected even where it is exposed.
- **Common alloys and their components**: *Brass* is copper with zinc (instruments, 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 Match the product to the conditions. Cold water with a reactive metal gives the *hydroxide*, for example $\mathrm{Ca(OH)_2}$; steam with a middle metal gives the *oxide*, for example ZnO. Writing the oxide for a cold-water reaction, or the hydroxide for a steam reaction, is a frequent slip.

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.

TIP The marking points are always the same three: the attached metal is *more reactive* than iron; it *loses electrons and corrodes in preference* to the iron; so the iron is *protected* while the sacrificial metal is used up. Connecting iron to a *less* reactive metal such as copper makes it rust faster.

TIP When explaining electrical conductivity, state that the *delocalised electrons* are free to move through the lattice and carry the charge. The positive ions stay fixed in the lattice. Saying that the ions move and carry the current is a common dropped mark.

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 pentane has $n = 5$, so its formula is C_5H_{12}. Alkanes are *saturated*, containing only single carbon-carbon bonds, which is why they hold the maximum possible number of hydrogen atoms.

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 hydrogens 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 every typical acid reaction.

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 by a halogen

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

Use for the *substitution* reaction in which one hydrogen of an alkane is replaced by one halogen atom, here forming chloromethane. The reaction needs *ultraviolet light* to start it and gives a hydrogen halide as the second product.

KEY CONCEPTS

- **Addition and condensation polymerisation**: 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, forming no other product. In *condensation polymerisation* two monomers, each with two reactive groups, join with the loss of a small molecule (usually water) at every linkage, forming a polyester or a polyamide.

- **Homologous series and functional groups**: A *functional group* is the atom or group of atoms that gives a molecule its characteristic reactions, such as the $-OH$ group of the alcohols. A *homologous series* is a family of compounds that share the same functional group and the same general formula, whose consecutive members differ by one CH_2 unit. Within a series the functional group fixes the chemical reactions while the chain length sets the physical properties, so boiling point rises steadily down the series.
- **How an organic name is built**: An organic name is built from a *stem* that counts the carbon atoms in the longest chain (meth, eth, prop, but) and a *suffix* that names the functional group (*-ane*, *-ene*, *-ol*, *-oic acid*). A *locant* number is inserted when the group could sit on more than one carbon, as in but-2-ene, and is omitted when only one position is possible, as in ethene.
- **Petroleum and fractional distillation**: *Petroleum* (crude oil) is a mixture of hydrocarbons, mostly alkanes 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. Down the column the chains are shorter: viscosity falls, the colour lightens and flammability rises.
- **Testing for unsaturation with bromine water**: An *unsaturated* compound contains at least one $C = C$ 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, whereas an alkane has no double bond and leaves the colour unchanged.
- **Addition reactions of the alkenes**: Alkenes are reactive because the $C = C$ double bond is a region of high electron density. In an *addition* reaction the double bond opens and a small molecule adds across it to give a single product: bromine forms a dibromoalkane, hydrogen with a nickel catalyst at about 150 °C forms the alkane, and steam with a phosphoric acid catalyst at about 300 °C and 60 atm forms an alcohol.
- **Carboxylic acids and esterification**: Carboxylic acids are *weak acids* that show the usual acid reactions, forming a salt (an *-oate*) with metals, bases and carbonates. Warming a carboxylic acid with an alcohol and a concentrated sulfuric acid catalyst gives an *ester* plus water, a reaction called *esterification*. In the ester name the alkyl part comes from the alcohol and the *-oate* part from the acid.
- **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 batch process giving impure ethanol that must be distilled. The *catalytic addition* of steam to ethene from petroleum is fast and continuous and gives pure ethanol directly, but uses a non-renewable raw material.

EXAM TIPS

TIP To find the monomer of an *addition* polymer, take one repeat unit, erase the continuation bonds at each end and put the $C = C$ double bond back between the two backbone carbons. So the repeat unit $-(CH_2 - CH_2)-$ comes from the monomer ethene, C_2H_4.

TIP Number the longest carbon chain from whichever end gives the functional group the *lowest* locant. For $CH_3CH = CHCH_3$ the double bond is reached at carbon 2 from either end, so it is but-2-ene. Quoting the higher number, or omitting the locant where two positions are possible, loses the mark.

10 States of matter

KEY FORMULAS

Relative molecular mass

$$M_r = \sum A_r$$

Use to find the relative molecular mass of a gas by adding the relative atomic mass of every atom shown in its formula. In this chapter M_r is the quantity that fixes how fast a gas diffuses, so compute it before comparing two gases.

Pressure and volume of a gas

$$pV = \text{constant}$$

Use for a fixed mass of gas at constant temperature, where pressure and volume are inversely proportional. Halving the volume doubles the pressure; compressing a gas to one third of its volume makes the pressure three times as large.

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 net movement 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. 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.
- **Gas pressure from particle collisions**: A gas exerts pressure because its particles constantly collide with the container walls, each collision giving a tiny outward push. Raising the temperature at fixed volume makes the particles move faster, so they strike the walls harder and more often and the pressure rises. Reducing the volume makes the particles reach the walls more often, which also raises the pressure.
- **Rate of diffusion depends on relative molecular mass**: At a fixed temperature the rate at which a gas diffuses is governed by its relative molecular mass: the smaller the M_r, the faster the gas diffuses. To compare two gases, work out each M_r, and the gas with the smaller value diffuses faster and travels further in the same time.
- **Why the temperature is 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 At the same temperature all gas molecules have the same average kinetic energy, so a lighter molecule must move faster. The smaller the relative molecular mass, the faster the gas diffuses. Always *calculate* each M_r rather than judging from the size of the formula, because carbon monoxide ($M_r = 28$) is heavier than methane ($M_r = 16$) despite looking small.

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 increases the volume at constant pressure, and reducing the pressure increases the volume at constant temperature. 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

$$N = n \times N_A$$

Use to find how many atoms, molecules or ions are present, where $N_A = 6.02 \times 10^{23}$ per mole. Read the substance carefully, because one mole of O_2 holds twice as many atoms as molecules.

Relative atomic mass from isotopic abundances

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

Use when an element exists as two or more isotopes and you are given the mass and percentage abundance of each. The result is a *weighted* mean, so it always lies closer to the more abundant isotope.

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.

Gas volume of a gas at r.t.p.

$$V = n \times 24$$

Use to convert between an amount of gas n in moles and its volume V in dm³ at room temperature and pressure, where one mole of any gas occupies 24 dm³. Valid only for a gas at r.t.p., never for a solid, liquid or solution.

Relative molecular and 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 $40 + 2(16) + 2(1) = 74$.

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 charges cancel. Combine the ions so the total positive charge equals the total negative charge, then reduce to lowest terms. A polyatomic ion that takes a subscript above one must be bracketed first, as in $Ca(NO_3)_2$.
- **Empirical and molecular formula**: The *molecular formula* gives the actual number of atoms of each element in one molecule, for example ethane C_2H_6. The *empirical formula* gives the simplest whole-

number ratio of those atoms, here CH_3. An ionic compound is always written as its empirical formula, such as $NaCl$, because a giant lattice has no discrete molecules.

- **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* N_A. The mole is the hub of stoichiometry because mass, number of particles, gas volume and concentration are each linked to one another only through the amount in moles.
- **The three-move recipe for reacting quantities**: Every reacting-quantity calculation follows one pattern: *get to moles* of what you are given, *cross the equation* by multiplying by the ratio of 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* 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 gallium(III) sulfate is $Ga_2(SO_4)_3$.

TIP In a chemical reaction atoms are only rearranged, so the total mass of products equals the total mass of reactants. An apparent loss or gain of mass means a gas has escaped or been absorbed. Because a fixed reaction scales as a whole, an unknown mass can often be found by simple proportion without using moles.

12 The Periodic Table

KEY FORMULAS

Group I metal reacting with water

$$2M + 2H_2O \rightarrow 2MOH + H_2$$

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

Halogen displacement reaction

$$X_2 + 2Y^- \rightarrow 2X^- + Y_2$$

Use when a halogen X_2 is added to a solution of a halide Y^-. The reaction proceeds only if X is more reactive than Y, so chlorine displaces bromide and iodide, bromine displaces only iodide, and iodine displaces neither.

Ion charge predicted from Group number

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

Use to predict the charge of a simple ion straight 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 V to VII gains electrons to give a negative charge equal to eight minus the Group number. Group IV elements rarely form simple ions.

Formula of a binary ionic compound by charge balancing

$$M^{a+} + X^{b-} \rightarrow M_bX_a$$

Use to build the formula of an ionic compound so the charges cancel: the subscript on each ion is the size of the other ion's charge, then reduce to the lowest whole-number ratio. For example Al^{3+} with O^{2-} gives Al_2O_3.

Variable oxidation state and formula deduction

$$Fe^{2+} \rightarrow FeCl_2, \qquad Fe^{3+} \rightarrow Fe_2O_3$$

Use when a transition metal name carries a Roman numeral, which gives the charge of its ion. Balance that charge against the other ion as usual: iron(II) with chloride gives $FeCl_2$, while iron(III) with oxide gives Fe_2O_3.

KEY CONCEPTS

- **Group I: the alkali metals**: The Group I elements (lithium, sodium, potassium and below) are the *alkali metals*: soft, low-density, low-melting metals, each with one outer-shell electron. Every Group I atom loses that single electron to form a 1+ ion, so the Group has one fixed valency. They are stored under oil because they react quickly with oxygen and water vapour in air.
- **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 Cl_2. Each atom has seven outer electrons and gains one more to form a 1− *halide* ion such as Cl^-, or shares a pair of electrons to bond covalently.
- **Metals, non-metals and the dividing line**: Metals occupy the large region on the left of the table and non-metals the smaller region to the upper right, separated by a staircase line. Position predicts the ion sign: a metal has few outer electrons and loses them to form a positive ion, while a non-metal has a nearly full shell and gains electrons to form a negative ion. An element on the line, such as silicon, is a *metalloid* with intermediate properties.
- **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 *inert*. Their atoms

already have a complete outer shell (eight electrons, or two for helium), so they have no tendency to gain, lose or share electrons.

- **Reading the Periodic Table: Groups and Periods**: A *Group* is a vertical column; for Groups I to VII every element has the same number of outer-shell electrons, which gives the Group similar chemical properties. A *Period* is a horizontal row, and the Period number equals the number of occupied electron shells. The elements are arranged in order of increasing proton number.
- **Transition elements: the defining properties**: The *transition elements* form the central block between Group II and Group III. Compared with the Group I metals they are hard, dense and high-melting, and they show three signatures: *variable oxidation states* (for example Fe^{2+} and Fe^{3+}), *coloured compounds*, and use as *catalysts*, such as iron in the Haber process.
- **The reactivity order of the halogens**: Reactivity *decreases* down Group VII, so the order is chlorine, then bromine, then iodine. A *displacement reaction* uses this order: a more reactive halogen takes the place of a less reactive one in its halide. Down the Group the outer shell is further from the nucleus and more shielded, so an incoming electron is attracted less strongly and gained less easily.

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 hydrogen is displaced from the water. A lilac flame seen with potassium is the escaping hydrogen igniting, not a new gas.

TIP The rule that the Group number equals the number of outer-shell electrons holds only for the main groups, I to VII. The central transition elements fill an inner sub-shell instead, so their column position does not give the outer-electron count, and they form ions of more than one charge rather than a single fixed valency.

TIP Other atoms react in order to reach a full, noble-gas outer shell; the noble gases already have one, so they have no reason to react. This is why a sodium atom loses one electron to gain neon's arrangement and a chlorine atom gains one to reach argon's.
