# Chemistry (0620)

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Compact 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.

### 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.

### 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.

## 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.

### 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.

## 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 $\Delta T$ 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.

### 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.

### 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.

# 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 $\mathrm{cm}^3$ 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.

## 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.

# 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.

## 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.

## 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.

# 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.

## 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.

## 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.

# 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.

# 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.

## 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.

# 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.

## 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 $-OH$ 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 $C = C$ (alkenes), $-OH$ (alcohols), $-COOH$ (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.

## 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.

### 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.

# 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 $40 + 2(16) + 2(1) = 74$.

## 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 $6.02 \times 10^{23}$ 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.

## 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.

# 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 hydrogen gas. The gas released is hydrogen, never oxygen.

Ion charge predicted from Group number

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

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.

## 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 $Cl_2$. Each atom has seven outer electrons and gains one more to form a 1- *halide* ion such as $Cl^-$.
- **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.