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IGCSE CHEMISTRY

COMPLETE STUDY HANDBOOK

2026 EDITION | Cambridge IGCSE (0620/0971)

Comprehensive Notes • Key Definitions • Exam Tips • Worked Examples

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Welcome from Eclassopedia

Welcome to the Eclassopedia IGCSE Chemistry Handbook 2026. This comprehensive guide has been carefully crafted by experienced educators and subject specialists to give you everything you need to excel in your Cambridge IGCSE Chemistry examinations.

Whether you are just beginning your IGCSE journey or in the final stages of revision, this handbook covers all syllabus topics for Cambridge IGCSE Chemistry (0620/0971) in a clear, structured, and student-friendly format.

> **This handbook aligns fully with the Cambridge IGCSE Chemistry Syllabus for examinations in 2026. Always cross-reference with your official Cambridge syllabus document for the most current assessment objectives.**

At Eclassopedia, our mission is to empower every learner with quality educational resources that are accessible, engaging, and effective. We believe that every student, regardless of background, deserves access to world-class study materials.

How to Use This Handbook

This handbook is organised to mirror the Cambridge IGCSE Chemistry syllabus. Each chapter corresponds to a syllabus topic and contains:

- Clear and concise topic notes written in accessible language
- Key definitions highlighted for quick revision
- Important equations and formulae presented in easy-to-read format
- Worked examples demonstrating how to apply concepts
- Exam tips and common pitfalls to avoid
- Summary tables for quick reference

We recommend reading through each chapter systematically during your studies, then using the key definitions and summary tables for rapid revision closer to your exams.

Chapter 1: The Particulate Nature of Matter

1.1 States of Matter

All matter exists in one of three states: solid, liquid, or gas. The state of a substance depends on the temperature and pressure, which affect the arrangement and movement of particles.

The kinetic particle theory explains the properties of each state by describing the behaviour of particles at a microscopic level.

State	Arrangement	Movement	Energy
Solid	Regular, fixed lattice	Vibrate about fixed positions	Lowest
Liquid	Random, close together	Move freely around each other	Intermediate
Gas	Random, widely spread	Move rapidly in all directions	Highest

1.2 Changes of State

When matter changes state, the arrangement and energy of particles change, but the chemical identity of the substance remains the same. Changes of state are physical changes.

- Melting: Solid to liquid (absorbs energy)
- Freezing: Liquid to solid (releases energy)
- Evaporation / Boiling: Liquid to gas (absorbs energy)
- Condensation: Gas to liquid (releases energy)
- Sublimation: Solid directly to gas (absorbs energy)
- Deposition: Gas directly to solid (releases energy)

> **Exam Tip: During a change of state, the temperature remains constant even though energy is being absorbed or released. This energy goes into breaking or forming intermolecular forces rather than increasing kinetic energy.**

1.3 Diffusion

Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration. It occurs in liquids and gases, but not in solids, because particles in solids cannot move freely.

Factors affecting the rate of diffusion:

- Temperature: Higher temperature increases the rate of diffusion because particles have greater kinetic energy
- Relative molecular mass (Mr): Lighter molecules diffuse faster than heavier ones
- Concentration gradient: A steeper gradient increases the rate

> **Key Formula: Graham's Law - Lighter gases diffuse faster. Rate of diffusion is inversely proportional to the square root of molecular mass.**

Chapter 2: Experimental Techniques

2.1 Criteria of Purity

A pure substance has a fixed, sharp melting point and boiling point. Impurities cause the melting point to be lower and the boiling point to be higher than the pure substance. They also cause a range of temperatures (called melting point depression) rather than a sharp transition.

> **Exam Tip: If a question asks how you can tell a substance is pure, state that it has a sharp melting point that matches the known value for the pure substance.**

2.2 Methods of Purification

Several laboratory techniques are used to separate and purify mixtures:

Filtration

Used to separate an insoluble solid from a liquid. The mixture is poured through filter paper in a funnel. The solid (residue) remains on the filter paper, while the liquid (filtrate) passes through.

Crystallisation

Used to obtain a pure solid from a solution. The solution is heated to evaporate some solvent, then cooled slowly. Crystals form as the solution becomes saturated. The crystals are filtered off and dried.

Simple Distillation

Used to obtain a pure liquid from a solution where the solute is non-volatile. The solution is heated and the vapour is condensed back into liquid by cooling. Used to obtain water from salt solution.

Fractional Distillation

Used to separate a mixture of liquids with different boiling points. The liquid with the lower boiling point distils over first. Used to separate crude oil into fractions or to separate ethanol from water.

Paper Chromatography

Used to separate and identify components in a mixture of dissolved substances. A spot of the mixture is placed on chromatography paper. A solvent moves up the paper by capillary action, carrying the components different distances depending on their solubility.

`Rf = distance moved by substance / distance moved by solvent front`

Each substance has a characteristic Rf value under given conditions. Rf values are used to identify unknown substances by comparison with known standards.

2.3 Identification of Gases

Gas	Test	Positive Result
Hydrogen (H_2)	Burning splint	Squeaky pop
Oxygen (O_2)	Glowing splint	Splint relights
Carbon dioxide (CO_2)	Limewater	Turns milky/cloudy
Chlorine (Cl_2)	Damp litmus paper	Bleaches to white
Ammonia (NH_3)	Damp red litmus paper	Turns blue

Chapter 3: Atoms, Elements and Compounds

3.1 Atomic Structure

An atom is the smallest particle of an element that can take part in a chemical reaction. Atoms consist of a small, dense nucleus surrounded by electrons.

Particle	Relative Mass	Relative Charge	Location
Proton	1	+1	Nucleus
Neutron	1	0	Nucleus
Electron	1/1840 (negligible)	-1	Orbitals around nucleus

> **Key Definitions: Atomic number (proton number) = number of protons in an atom. Mass number = number of protons + neutrons. For a neutral atom, number of electrons = number of protons.**

3.2 Isotopes

Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons. They have the same chemical properties (because they have the same electron configuration) but different physical properties (because they have different masses).

Examples of isotopes:

- Carbon-12 and Carbon-14 (both have 6 protons; C-12 has 6 neutrons, C-14 has 8 neutrons)
- Chlorine-35 and Chlorine-37 (both have 17 protons)
- Hydrogen (H-1), Deuterium (H-2), Tritium (H-3)

3.3 Electronic Configuration

Electrons occupy energy levels (shells) around the nucleus. The first shell holds a maximum of 2 electrons; the second and third shells hold a maximum of 8 electrons each.

Electronic configurations of the first 20 elements:

Element	Symbol	Atomic No.	Configuration
Hydrogen	H	1	1
Helium	He	2	2

Element	Symbol	Atomic No.	Configuration
Lithium	Li	3	2, 1
Carbon	C	6	2, 4
Nitrogen	N	7	2, 5
Oxygen	O	8	2, 6
Neon	Ne	10	2, 8
Sodium	Na	11	2, 8, 1
Chlorine	Cl	17	2, 8, 7
Argon	Ar	18	2, 8, 8
Calcium	Ca	20	2, 8, 8, 2

Chapter 4: Stoichiometry

4.1 The Mole Concept

The mole is the SI unit for the amount of substance. One mole of any substance contains the same number of particles as there are atoms in exactly 12 g of carbon-12. This number is the Avogadro constant.

`Avogadro's Constant (L) = 6.02 x 10^23 mol^-1`

`Number of moles = Mass (g) / Relative Atomic (or Molecular) Mass`

4.2 Molar Mass

The molar mass of a substance is the mass of one mole of that substance, expressed in g/mol. It is numerically equal to the relative molecular mass (Mr) of the substance.

To calculate Mr, add together the relative atomic masses (Ar) of all atoms in the formula:

- H_2O: Mr = 2(1) + 16 = 18 g/mol
- CO_2: Mr = 12 + 2(16) = 44 g/mol
- NaCl: Mr = 23 + 35.5 = 58.5 g/mol
- H_2SO_4: Mr = 2(1) + 32 + 4(16) = 98 g/mol

4.3 Reacting Masses

Balanced chemical equations can be used to calculate the masses of reactants and products in a reaction. The ratio of moles in a balanced equation is the mole ratio.

> **Steps for reacting mass calculations: (1) Write the balanced equation. (2) Write the mole ratio under the relevant species. (3) Convert moles to mass using molar mass. (4) Scale to the given quantity.**

Worked Example:

What mass of water is produced when 4 g of hydrogen burns in excess oxygen?

$2H_2 + O_2 \rightarrow 2H_2O$

Moles of H_2 = 4 / 2 = 2 mol

From equation: 2 mol H_2 produces 2 mol H_2O

Mass of H_2O = 2 x 18 = 36 g

4.4 Empirical and Molecular Formulae

The empirical formula shows the simplest whole-number ratio of atoms in a compound. The molecular formula shows the actual number of atoms in one molecule.

To find empirical formula from percentage composition:

- Divide each percentage by the element's Ar to get moles
- Divide by the smallest number of moles to get the ratio
- If necessary, multiply to get whole numbers

`Molecular formula = n x Empirical formula   (where n = Mr / empirical formula mass)`

4.5 Molar Volume of Gases

At standard temperature and pressure (STP: 0°C, 1 atm), one mole of any gas occupies 22.4 dm3. At room temperature and pressure (RTP: 25°C, 1 atm), one mole of gas occupies approximately 24 dm3.

`Volume of gas (dm3) = Moles x Molar Volume (24 dm3/mol at RTP)`

Chapter 5: Electricity and Chemistry

5.1 Electrolysis

Electrolysis is the decomposition of a substance using electricity. It requires an electrolyte (a molten ionic compound or an aqueous solution of an ionic compound), two electrodes (cathode and anode), and a direct current (DC) power supply.

- Cathode: Negative electrode — attracts cations (positive ions), which gain electrons (reduction occurs)
- Anode: Positive electrode — attracts anions (negative ions), which lose electrons (oxidation occurs)

> **Memory Aid: Oil Rig — Oxidation Is Loss, Reduction Is Gain (of electrons). Oxidation occurs at the Anode; Reduction occurs at the Cathode.**

5.2 Products of Electrolysis

The products of electrolysis depend on the electrolyte used and whether it is molten or aqueous. For aqueous solutions, the relative concentrations of ions affect the products at each electrode.

Electrolyte	Product at Cathode	Product at Anode
Molten lead(II) bromide	Lead (Pb)	Bromine (Br2)
Dilute sulfuric acid	Hydrogen (H2)	Oxygen (O2)
Dilute sodium chloride	Hydrogen (H2)	Oxygen (O2)
Concentrated sodium chloride	Hydrogen (H2)	Chlorine (Cl2)
Copper(II) sulfate (Cu electrodes)	Copper (Cu)	Copper dissolves

5.3 Industrial Applications of Electrolysis

Electrolysis has several important industrial applications:

Electroplating

A thin layer of metal is deposited on an object to protect it from corrosion or improve its appearance. The object is made the cathode. The metal to be plated is used as the anode.

Extraction of Aluminium

Aluminium is extracted from molten aluminium oxide (bauxite) by electrolysis. The high melting point of aluminium oxide means the process requires large amounts of energy, making aluminium expensive to produce.

Purification of Copper

Impure copper is refined by electrolysis. Impure copper acts as the anode, pure copper as the cathode, and copper(II) sulfate solution as the electrolyte. Copper dissolves from the anode and deposits at the cathode.

Chapter 6: Chemical Energetics

6.1 Exothermic and Endothermic Reactions

Chemical reactions involve the transfer of energy, usually as heat. Reactions can be classified as exothermic or endothermic based on the direction of energy transfer.

Type	Energy Transfer	Temperature Change	Examples
Exothermic	Released to surroundings	Temperature increases	Combustion, neutralisation, respiration
Endothermic	Absorbed from surroundings	Temperature decreases	Thermal decomposition, photosynthesis

6.2 Enthalpy Change (ΔH)

The enthalpy change (ΔH) measures the heat energy exchanged at constant pressure. It is measured in kJ/mol.

- Exothermic reactions: ΔH is negative (energy is released)
- Endothermic reactions: ΔH is positive (energy is absorbed)

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q = mcΔT   (where q = heat change in J, m = mass in g, c = specific
            heat capacity, ΔT = change in temperature)
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Worked Example:

50 cm3 of 1.0 mol/dm3 HCl is mixed with 50 cm3 of 1.0 mol/dm3 NaOH. The temperature rises by 6.8°C. Calculate the enthalpy change of neutralisation.

Total mass = 100 g (assuming density = 1 g/cm3)

q = 100 x 4.18 x 6.8 = 2842 J = 2.84 kJ

Moles of water formed = 0.05 mol

ΔH = -2.84 / 0.05 = -56.8 kJ/mol

6.3 Bond Energies

During reactions, chemical bonds are broken and formed. Breaking bonds requires energy (endothermic); forming bonds releases energy (exothermic). The overall enthalpy change depends on the difference between energy absorbed and energy released.

`ΔH = Sum of bond energies broken - Sum of bond energies formed`

> **Exam Tip: If bonds broken > bonds formed (in terms of energy), ΔH is positive (endothermic). If bonds formed > bonds broken, ΔH is negative (exothermic).**

Chapter 7: Chemical Reactions

7.1 Rate of Reaction

The rate of reaction is the change in concentration of a reactant or product per unit time. Reactions occur when particles collide with sufficient energy (activation energy) and the correct orientation.

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Rate of reaction = Change in amount of reactant or product / Time
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7.2 Factors Affecting Rate of Reaction

The following factors affect the rate of a chemical reaction:

Concentration

Increasing the concentration of a reactant increases the rate of reaction because there are more particles per unit volume, leading to more frequent collisions.

Temperature

Increasing the temperature increases the rate of reaction. Particles move faster, so collisions are more frequent. More importantly, a greater proportion of particles have energy greater than or equal to the activation energy, so more collisions result in reaction.

Surface Area

Increasing the surface area of a solid reactant increases the rate of reaction. More surface area means more particles are exposed and available for collisions with reactant particles.

Pressure (gases only)

Increasing the pressure of gaseous reactants increases the rate of reaction because the particles are more concentrated, leading to more frequent collisions.

Catalysts

A catalyst increases the rate of reaction by providing an alternative reaction pathway with a lower activation energy. The catalyst is not consumed in the reaction and can be recovered unchanged.

> **Exam Tip: When explaining the effect of any factor, you must refer to collision frequency AND the proportion of successful collisions for full marks.**

7.3 Reversible Reactions and Equilibrium

A reversible reaction is one that can proceed in both directions. At equilibrium, the rate of the forward reaction equals the rate of the reverse reaction, and the concentrations of reactants and products remain constant.

Le Chatelier's Principle states that if a system at equilibrium is subjected to a change, the system will adjust to oppose that change and establish a new equilibrium.

Change Applied	Effect on Equilibrium
Increase concentration of reactant	Equilibrium shifts to the right (forward reaction favoured)
Decrease concentration of product	Equilibrium shifts to the right
Increase temperature	Equilibrium shifts in the endothermic direction
Increase pressure	Equilibrium shifts to the side with fewer moles of gas
Add a catalyst	No effect on position of equilibrium; equilibrium is reached faster

Chapter 8: Acids, Bases and Salts

8.1 Properties of Acids

An acid is a substance that donates protons (H+ ions) in aqueous solution. Acids have pH values below 7. Common laboratory acids include hydrochloric acid (HCl), sulfuric acid (H_2SO_4), and nitric acid (HNO_3).

Properties of acids in aqueous solution:

- Turn blue litmus red
- Have pH less than 7
- React with metals to produce hydrogen gas
- React with carbonates to produce carbon dioxide
- React with metal oxides and hydroxides (bases) in neutralisation reactions

8.2 Properties of Bases and Alkalis

A base is a substance that accepts protons (H+ ions). An alkali is a base that is soluble in water and produces hydroxide ions (OH-) in solution. Alkalis have pH values above 7.

> **Key Distinction: All alkalis are bases, but not all bases are alkalis. Metal hydroxides that are insoluble (e.g. copper(II) hydroxide) are bases but not alkalis.**

8.3 The pH Scale

The pH scale runs from 0 to 14 and measures the acidity or alkalinity of a solution. pH 7 is neutral. Each unit on the pH scale represents a tenfold change in hydrogen ion concentration.

- pH 0-6: Acidic (lower pH = stronger acid = higher [H+])
- pH 7: Neutral
- pH 8-14: Alkaline (higher pH = stronger alkali = lower [H+])

8.4 Preparation of Salts

Salts can be prepared by several methods depending on their solubility:

Soluble Salts

- Acid + Metal: e.g. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
- Acid + Metal Oxide: e.g. $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

- Acid + Metal Hydroxide (Neutralisation): e.g. $NaOH + HCl \rightarrow NaCl + H_2O$
- Acid + Metal Carbonate: e.g. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

Insoluble Salts

Insoluble salts are prepared by precipitation - mixing two solutions of soluble compounds that contain the required ions.

Example: $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

8.5 Solubility Rules

Ion	Solubility	Exceptions
Nitrates (NO_3^-)	All soluble	None
Chlorides (Cl^-)	Most soluble	$AgCl$, $PbCl_2$ insoluble
Sulfates (SO_4^{2-})	Most soluble	$BaSO_4$, $PbSO_4$, $CaSO_4$ insoluble
Carbonates (CO_3^{2-})	Most insoluble	Na_2CO_3, K_2CO_3, $(NH_4)_2CO_3$ soluble
Hydroxides (OH^-)	Most insoluble	$NaOH$, KOH, $Ca(OH)_2$ soluble

Chapter 9: The Periodic Table

9.1 Organisation of the Periodic Table

The Periodic Table arranges elements in order of increasing atomic number. Elements with similar properties are placed in the same vertical column (group). Elements in the same horizontal row are in the same period.

- Groups (vertical columns, 1-18): Elements in the same group have the same number of electrons in their outer shell, giving similar chemical properties
- Periods (horizontal rows, 1-7): Elements in the same period have the same number of electron shells
- Metals are on the left and centre; non-metals are on the right
- The zigzag line separates metals from non-metals; semi-metals (metalloids) lie along this line

9.2 Group 1 - Alkali Metals

The alkali metals (Li, Na, K, Rb, Cs, Fr) are soft metals that react vigorously with water and oxygen. Reactivity increases down the group as the outer electron is further from the nucleus and more easily lost.

$$2Na + 2H2O \rightarrow 2NaOH + H2$$

Trends down Group 1:

- Increasing reactivity
- Decreasing melting and boiling points
- Increasing atomic radius
- All form ions with a 1+ charge (e.g. Li+, Na+, K+)

9.3 Group 7 - Halogens

The halogens (F, Cl, Br, I, At) are non-metals that exist as diatomic molecules (F2, Cl2, Br2, I2). Reactivity decreases down the group as it becomes harder to gain an electron.

Halogen	State at Room Temp.	Colour	Reactivity
Fluorine (F2)	Gas	Pale yellow	Most reactive
Chlorine (Cl2)	Gas	Yellow-green	Very reactive

Halogen	State at Room Temp.	Colour	Reactivity
Bromine (Br_2)	Liquid	Red-brown	Moderately reactive
Iodine (I_2)	Solid	Grey-black (purple vapour)	Less reactive

Displacement reactions: A more reactive halogen will displace a less reactive halogen from its aqueous salt solution.

$$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$$

9.4 Group 0 - Noble Gases

The noble gases (He, Ne, Ar, Kr, Xe, Rn) are colourless, odourless, monatomic gases with very low reactivity. Their full outer electron shells (8 electrons, except He with 2) make them extremely stable.

Uses of noble gases:

- Helium: filling balloons and airships (lighter than air, non-flammable)
- Argon: filling light bulbs to prevent filament burning
- Neon: advertising signs (glows red/orange when electricity passes through)

9.5 Transition Metals

Transition metals occupy the central block of the periodic table (periods 4-7). They have several characteristic properties:

- High melting points and densities
- Good conductors of heat and electricity
- Variable oxidation states (e.g. iron can be Fe^{2+} or Fe^{3+})
- Form coloured compounds (e.g. copper compounds are blue, iron(II) compounds are green)
- Act as catalysts (e.g. iron in Haber process, platinum in catalytic converters)

Chapter 10: Metals

10.1 Properties of Metals

Metals have a giant metallic structure in which positive metal ions are surrounded by a sea of delocalised electrons. This structure explains the characteristic properties of metals.

Property	Explanation
Good conductors of electricity	Delocalised electrons can move through the structure and carry charge
Good conductors of heat	Delocalised electrons and vibrating ions transfer heat energy
Malleable and ductile	Layers of ions can slide over each other without breaking the metallic bond
High melting/boiling points	Strong electrostatic attraction between positive ions and electron sea
Lustrous (shiny)	Delocalised electrons reflect light

10.2 Reactivity Series

The reactivity series ranks metals in order of their reactivity. More reactive metals displace less reactive metals from their compounds in solution or when heated.

Reactivity Series (most to least reactive):

- Potassium (K) — most reactive
- Sodium (Na)
- Calcium (Ca)
- Magnesium (Mg)
- Aluminium (Al)
- Zinc (Zn)
- Iron (Fe)
- Lead (Pb)
- Copper (Cu)
- Silver (Ag)
- Gold (Au) — least reactive

> **Memory Aid: Please Stop Calling Me A Zebra In London, Can She Go. (K, Na, Ca, Mg, Al, Zn, Fe, Pb, Cu, Ag, Au)**

10.3 Extraction of Metals

The method used to extract a metal depends on its position in the reactivity series:

- Very reactive metals (K, Na, Ca, Mg, Al): Extracted by electrolysis of molten compounds
- Moderately reactive metals (Zn, Fe, Pb): Extracted by reduction with carbon (smelting)
- Unreactive metals (Cu, Ag, Au): Found native (uncombined) or extracted by simple chemical processes

Extraction of Iron (Blast Furnace)

Iron is extracted from iron ore (haematite, $Fe2O3$) in a blast furnace using coke (carbon) as the reducing agent. Limestone ($CaCO3$) is added to remove acidic impurities (slag).

$$Fe2O3 + 3CO \rightarrow 2Fe + 3CO2$$

10.4 Corrosion

Corrosion is the degradation of a metal by a chemical reaction with its environment. Rusting is the specific corrosion of iron and steel, forming hydrated iron(III) oxide (rust). Both water and oxygen are required for rusting.

Methods of preventing corrosion:

- Painting or coating with oil/grease: Barrier protection
- Electroplating: Coating with a less reactive metal
- Galvanising: Coating with zinc (also provides sacrificial protection)
- Sacrificial protection: Attaching a more reactive metal (zinc or magnesium) that corrodes preferentially
- Alloying: Mixing with other metals to reduce reactivity (e.g. stainless steel contains chromium)

Chapter 11: Organic Chemistry

11.1 Introduction to Organic Chemistry

Organic chemistry is the study of carbon-containing compounds. Carbon atoms have 4 valence electrons and can form 4 covalent bonds, making them capable of forming an enormous variety of molecules. Organic compounds are grouped into homologous series — families of compounds with the same functional group and general formula, showing a gradual change in properties.

11.2 Alkanes

Alkanes are saturated hydrocarbons (contain only single C-C bonds). They have the general formula C_nH_{2n+2}.

Name	Formula	Structural Formula	State at RTP
Methane	CH_4	CH_4	Gas
Ethane	C_2H_6	CH_3CH_3	Gas
Propane	C_3H_8	$CH_3CH_2CH_3$	Gas
Butane	C_4H_{10}	$CH_3(CH_2)_2CH_3$	Gas
Pentane	C_5H_{12}	$CH_3(CH_2)_3CH_3$	Liquid

Alkanes undergo combustion and substitution reactions (with halogens in the presence of UV light).

Complete combustion: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

11.3 Alkenes

Alkenes are unsaturated hydrocarbons that contain at least one C=C double bond. They have the general formula C_nH_{2n}. The presence of a double bond makes alkenes more reactive than alkanes.

Alkenes undergo addition reactions, in which a molecule adds across the C=C double bond:

- Addition of hydrogen (hydrogenation): $C_2H_4 + H_2 \rightarrow C_2H_6$ (requires Ni catalyst, 150°C)
- Addition of bromine (bromination): $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$ (decolourises bromine water — test for C=C)
- Addition of water (hydration): $C_2H_4 + H_2O \rightarrow C_2H_5OH$ (requires H_3PO_4 catalyst, 300°C, 60 atm)

> **Test for Alkenes: Alkenes decolourise bromine water (from orange-brown to colourless). Alkanes do not react with bromine water.**

11.4 Alcohols

Alcohols contain the hydroxyl functional group (-OH). The general formula for simple alcohols is $C_nH_{2n+1}OH$. Ethanol (C_2H_5OH) is the most important commercially.

Properties and reactions of alcohols:

- Combustion: Alcohols burn in air producing CO_2 and H_2O
- Oxidation: Alcohols are oxidised to carboxylic acids (using acidified potassium dichromate as oxidising agent)
- Dehydration: Alcohols lose water to form alkenes when heated with a concentrated acid catalyst
- Fermentation: Glucose is converted to ethanol and CO_2 by yeast enzymes

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2 \quad \text{(fermentation of glucose)}$$

11.5 Carboxylic Acids

Carboxylic acids contain the carboxyl functional group (-COOH). Ethanoic acid (CH_3COOH) is found in vinegar. Carboxylic acids are weak acids — they partially dissociate in water.

Reactions of carboxylic acids:

- Neutralisation with bases to form salts and water
- Reaction with carbonates to produce a salt, water, and CO_2
- Esterification: React with alcohols to form esters

$$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O \quad \text{(esterification, reversible)}$$

11.6 Polymers

Polymers are large molecules formed by the joining together of many small molecules called monomers.

Addition Polymerisation

Alkene monomers join together via their double bonds to form addition polymers. No other products are formed.

- Polyethene (polythene): from ethene monomers — used in plastic bags, packaging
- Polypropylene: from propene monomers — used in ropes, carpet

- PVC: from chloroethene monomers — used in pipes, window frames

Condensation Polymerisation

Monomers with two functional groups join together, releasing small molecules (usually water) as a by-product.

- Polyesters (e.g. Terylene/PET): formed from a diol and a dicarboxylic acid, releasing water
- Polyamides (e.g. Nylon): formed from a diamine and a dicarboxylic acid, releasing water

> **Exam Tip: To identify a condensation polymer, look for loss of water (or HCl). To identify an addition polymer, look for C=C bonds in the monomer and no small molecule by-product.**

Chapter 12: Environmental Chemistry

12.1 Air and Atmosphere

The atmosphere consists mainly of nitrogen (78%) and oxygen (21%), with small amounts of argon (0.9%), carbon dioxide (0.04%), and water vapour. This composition has remained relatively stable for millions of years through natural cycles.

12.2 Air Pollution

Human activity, particularly the burning of fossil fuels, releases pollutant gases into the atmosphere. These pollutants have serious effects on the environment and human health.

Pollutant	Source	Effect	Solution
Carbon monoxide (CO)	Incomplete combustion	Toxic; binds to haemoglobin	Catalytic converters
Sulfur dioxide (SO2)	Burning fossil fuels (sulfur impurities)	Acid rain; respiratory problems	Remove sulfur from fuels; FGD
Nitrogen oxides (NOx)	High-temperature combustion in engines	Acid rain; photochemical smog	Catalytic converters; EGR
Particulates	Burning fuels; industry	Respiratory disease; reduced visibility	Filters; clean fuel
Carbon dioxide (CO2)	Complete combustion; deforestation	Enhanced greenhouse effect	Renewable energy; reforestation

12.3 The Greenhouse Effect

The greenhouse effect is a natural process that keeps Earth warm. Greenhouse gases (CO2, CH4, H2O, N2O) absorb infrared radiation emitted by Earth's surface and re-radiate it back, preventing heat from escaping to space.

Human activities have increased greenhouse gas concentrations (particularly CO2 and CH4), enhancing the greenhouse effect and causing global climate change. Consequences include rising temperatures, melting ice caps, rising sea levels, and more extreme weather events.

> **Key Distinction: The greenhouse effect is natural and essential for life. The enhanced greenhouse effect (due to human activity) causes climate change.**

12.4 Water Chemistry

Water is essential for all life and is widely used in industry and households. Water treatment makes fresh water safe for drinking.

Water treatment stages:

- Sedimentation: Allowing large particles to settle
- Flocculation: Adding aluminium sulfate to cause small particles to clump together
- Filtration: Passing through sand and gravel to remove suspended particles
- Chlorination: Adding chlorine gas or compounds to kill bacteria and pathogens

Distillation can produce pure water but is expensive and energy-intensive. Ion exchange resins can remove dissolved ions to produce deionised water.

Chapter 13: Exam Preparation and Revision

13.1 Overview of Assessment

Cambridge IGCSE Chemistry (0620/0971) is assessed through a combination of written papers and practical work. Examinations take place in May/June 2026.

Paper	Type	Duration	Marks	Weightage
Paper 1	Multiple choice	45 min	40	30%
Paper 2	Core theory	1h 15min	80	50%
Paper 4	Extended theory	1h 15min	80	50%
Paper 5	Practical test	1h 15min	40	20%
Paper 6	Alternative to practical	1h	40	20%

13.2 Key Revision Strategies

- Use active recall: Test yourself without looking at notes
- Create mind maps connecting topics (e.g. bonding links to properties, reactions, and uses)
- Practise past papers under timed conditions
- Learn definitions precisely — examiners often award marks for specific key words
- Understand rather than memorise: if you understand why, you can work out the answer
- Focus on command words: describe, explain, suggest, compare, state — each requires a different type of answer

13.3 Common Command Words

Command Word	What It Requires
State / Name / Give	A brief, factual answer with no explanation required
Describe	A detailed account of what happens (observations, not explanations)
Explain	Say why something happens, using chemical principles
Suggest	Use your knowledge to come up with a plausible idea
Compare	Identify similarities AND differences between two things
Calculate	Work out a numerical answer, showing your working
Deduce / Predict	Use given information or patterns to arrive at a conclusion

13.4 Top Exam Tips from Eclassopedia

- Always show your working in calculation questions — you can earn method marks even if your final answer is wrong
- Reread the question before answering — check the number of marks available as a guide to how much to write
- Use correct chemical vocabulary and avoid vague language
- For 'explain' questions about rate, mention collision frequency and activation energy
- Draw dot-and-cross diagrams carefully; show all outer electrons for both bonding pairs and lone pairs
- For ionic equations, check that charges and atoms balance on both sides
- In practical questions, suggest ways to make a method more accurate or reliable
- Review your periodic table — Ar values are given, but knowing common values speeds up calculations

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Quick Reference: Key Formulae and Constants

Formula / Constant	Expression	Units
Moles from mass	$n = m / Mr$	mol
Moles from volume of gas	$n = V / 24$	mol (at RTP)
Moles from solution	$n = c \times V$	mol
Concentration	$c = n / V$	mol/dm3
Enthalpy change from calorimetry	$q = mc\Delta T$	J or kJ
Enthalpy change per mole	$\Delta H = -q / n$	kJ/mol
Rf value (chromatography)	Rf = d(substance) / d(solvent)	No units (0-1)
Avogadro's constant	$L = 6.02 \times 10^{23}$	mol-1
Molar volume at RTP	$Vm = 24$	dm3/mol

Best of luck in your 2026 IGCSE Chemistry examinations!

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