CURRICULUM-ALIGNED LEARNING RESOURCES

GRADE 11 CHEMISTRY

COMPREHENSIVE SHORT NOTES

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STRAND 1.0	Inorganic Chemistry
STRAND 2.0	Physical Chemistry
STRAND 3.0	Organic Chemistry

A revision-focused guide for Kenya Senior School learners

About These Notes

These Comprehensive Short Notes are designed to support Grade 11 Chemistry learners following the Kenya Institute of Curriculum Development (KICD) Senior School Chemistry Curriculum Design. The notes preserve the curriculum's strand and sub-strand structure while explaining key ideas in a connected, revision-friendly way. They combine definitions, explanations, practical guidance, worked examples, tables, and opportunities for application.

The curriculum design identifies Strand 1.0: Inorganic Chemistry, with Sub-strand 1.1: The Mole and Sub-strand 1.2: Non-metals. The mole section focuses on amount of substance, relative masses, conversion between mass and moles, molar solutions, solution preparation and concentration, titration, redox titration, mole ratios, and molar gas volumes. The non-metals section focuses on oxygen and nitrogen, gas collection, laboratory preparation, industrial separation of air, physical and chemical properties, and uses.

How to Use the Notes

- Read each explanation and note the meaning of the symbols and units before attempting calculations.
- Work through each worked example in order; cover the solution and try the question independently before checking it.
- Use the highlighted yellow illustration prompts as placeholders. Generate the requested textbook-style diagram and replace the entire highlighted prompt in Word.
- Practise writing balanced chemical equations, stating observations, and explaining why a method or apparatus is appropriate.
- At the end of each major section, use the summary and self-check questions to identify concepts that need more revision.

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Strand Map

Strand / Sub-strand	Main content
1.0 Inorganic Chemistry	Quantitative chemistry and selected non-metals
1.1 The Mole	Relative masses; mole and Avogadro constant; mass–mole–particle calculations; molar solutions; concentration; titration; redox titration; mole ratios; molar gas volume
1.2 Non-metals	Oxygen and nitrogen; gas collection; laboratory preparation; industrial isolation; physical and chemical properties; uses of the gases and their compounds

Safety note: Laboratory work must be carried out under the supervision of a qualified teacher, using suitable protective equipment and the school's approved risk-control procedures. Never taste chemicals or directly inhale gases. Wafting is only appropriate when the teacher has explicitly authorised it and the substance is known to be suitable for that technique.

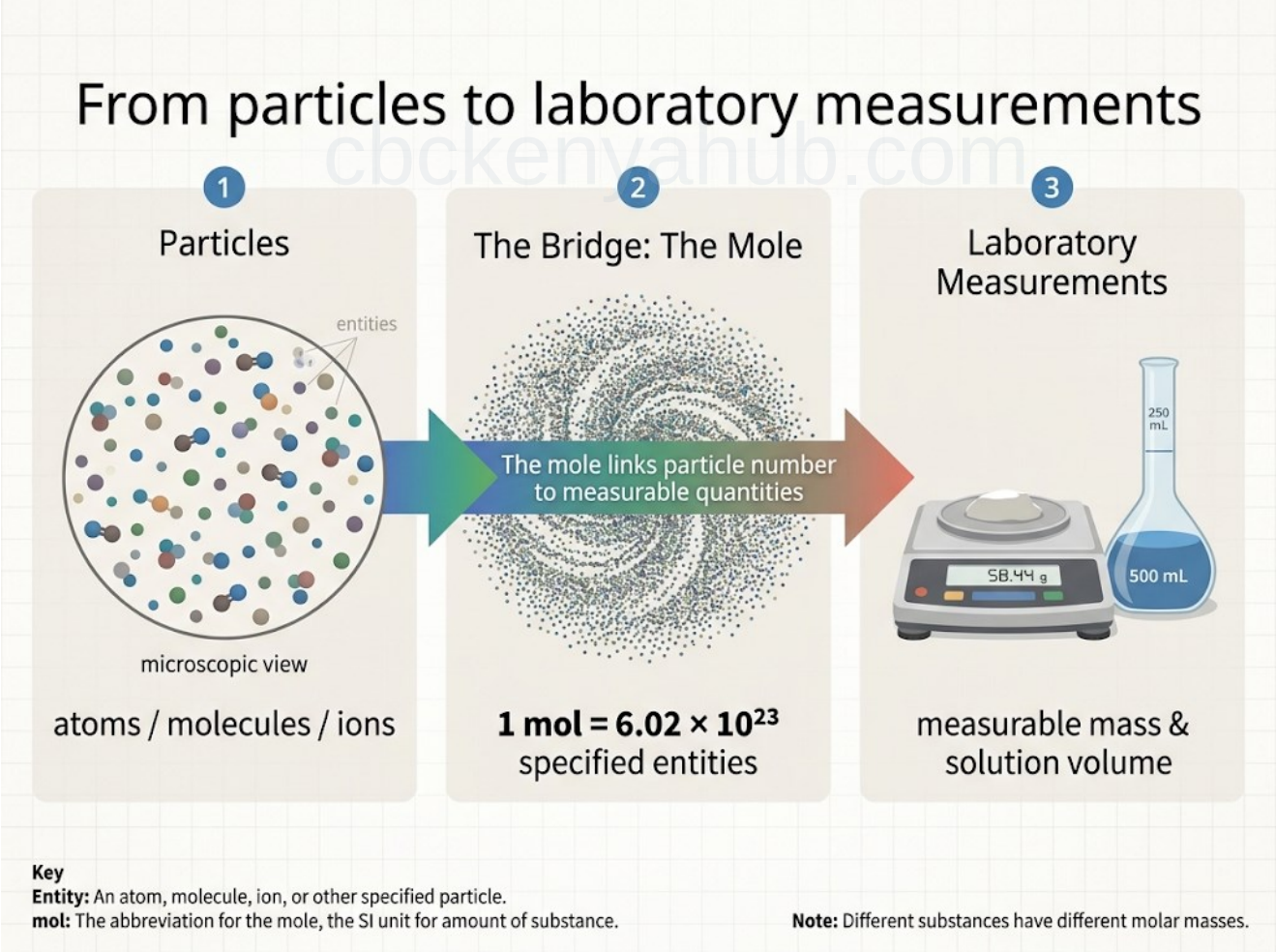
SUB-STRAND 1.1: THE MOLE

1.1.1 Why Chemists Need the Mole

Chemical reactions involve atoms, ions and molecules, but these particles are far too small to count individually in an ordinary laboratory. Chemists therefore measure a convenient bulk quantity and relate it to the number of particles present. The mole provides this bridge between the microscopic world of particles and measurable quantities such as mass, volume and concentration.

For example, a balanced equation can tell us that two molecules of hydrogen react with one molecule of oxygen to form two molecules of water. In the laboratory, however, we do not usually count individual molecules. We weigh substances and measure solutions. The mole allows the numerical ratios in a balanced equation to be used with these practical measurements.

The SI unit for amount of substance is the mole, symbol mol. One mole contains exactly $6.02214076 \times 10^{23}$ specified elementary entities. At Grade 11 level, this value is commonly rounded to 6.02×10^{23}. The entities must be identified: they may be atoms, molecules, ions, electrons or formula units, depending on the substance.



1.1.2 The Avogadro Constant

The Avogadro constant, N_a, is the number of specified entities in one mole. Its commonly used value is 6.02×10^{23} mol^{-1}. The superscript −1 means “per mole”. One mole of carbon atoms and one mole of oxygen molecules contain the same number of entities only when the entity is specified consistently; the mass of the two samples is different because carbon atoms and oxygen molecules have different masses.

A useful everyday analogy is a dozen. A dozen eggs and a dozen oranges contain the same number of items—12—but their total masses differ. Similarly, one mole of any specified entities contains the same number of entities, while its mass depends on the identity and formula of the substance.

Worked Example 1: Number of particles in 0.50 mol

Question: Calculate the number of molecules in 0.50 mol of oxygen gas, O_2.

Step 1: Write the relationship. Number of entities = amount in moles × Avogadro constant.

Step 2: Substitute. Number of O_2 molecules = 0.50 mol × 6.02×10^{23} mol^{-1}.

Step 3: Calculate and state the entity. Number of O_2 molecules = 3.01×10^{23} molecules.

Notice that the unit mol cancels with mol^{-1}. Always identify whether the final answer is a number of atoms, molecules, ions or formula units.

Worked Example 2: Finding amount from particle number

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Question: A sample contains 1.204×10^{24} sodium ions. Find the amount of sodium ions in moles.

Amount, n = number of entities ÷ N_a = $(1.204 \times 10^{24}) \div (6.02 \times 10^{23})$ = 2.00 mol (to three significant figures).

The answer is 2.00 mol of Na^+ ions, not 2.00 mol of sodium metal. The charge and identity of the entity matter.

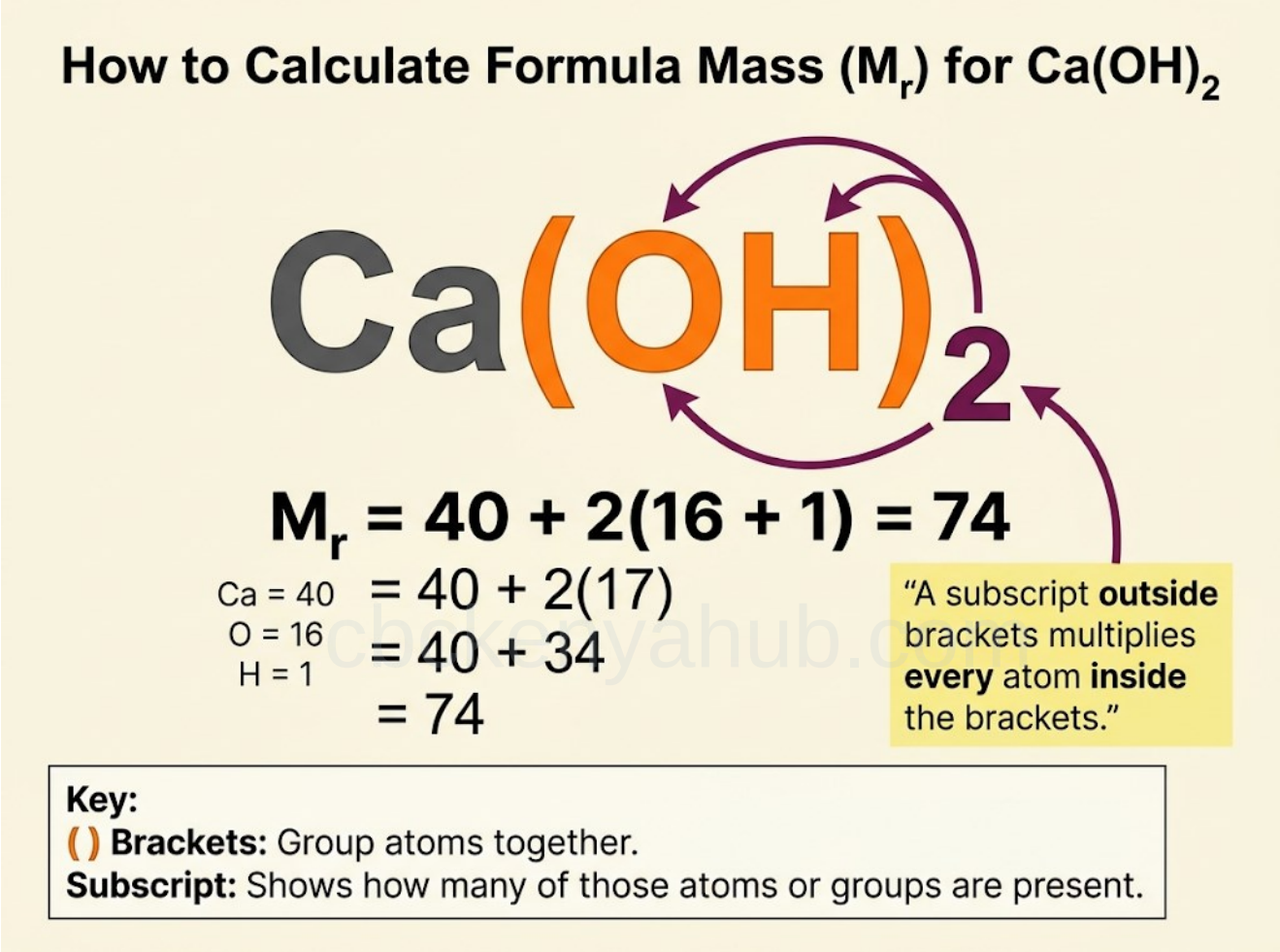
1.1.3 Relative Atomic Mass and Relative Formula Mass

Atoms of different elements have different masses. Since an individual atom is extremely light, chemists use relative atomic mass, A_r, to compare the average mass of atoms of an element with one-twelfth of the mass of a carbon-12 atom. Relative atomic mass has no unit because it is a ratio.

The relative atomic mass of an element is usually obtained from the weighted average of the masses of its naturally occurring isotopes. For routine Grade 11 calculations, use the atomic masses supplied in a question or the periodic table provided by the teacher. Common approximate values are H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, S = 32, Cl = 35.5, K = 39, Ca = 40 and Fe = 56.

Relative formula mass, M_r, is calculated by adding the relative atomic masses of all atoms represented in a chemical formula. For a covalent substance that exists as molecules, the same sum is often called relative molecular mass. Remember to multiply each atomic mass by the number of that atom shown by the subscript.

Formula	Calculation using common values	Relative formula / molecular mass
H_2O	$(2 \times 1) + 16$	18
CO_2	$12 + (2 \times 16)$	44
NH_3	$14 + (3 \times 1)$	17
$Ca(OH)_2$	$40 + 2(16 + 1)$	74
$Al_2(SO_4)_3$	$(2 \times 27) + 3[32 + (4 \times 16)]$	342



Worked Example 3: Relative formula mass of ammonium sulfate

Question: Calculate M_r for $(NH_4)_2SO_4$ using N = 14, H = 1, S = 32 and O = 16.

The formula contains 2 nitrogen atoms, 8 hydrogen atoms, 1 sulfur atom and 4 oxygen atoms. Therefore $M_r = (2 \times 14) + (8 \times 1) + 32 + (4 \times 16) = 28 + 8 + 32 + 64 = 132$.

A frequent error is to count only one NH_4 group. The subscript outside the brackets doubles both N and H in the ammonium ion.

1.1.4 Molar Mass

Molar mass, M, is the mass of one mole of a substance. It is expressed in g mol^{-1}. Its numerical value is equal to the relative atomic mass for an element or the relative formula mass for a compound when the usual atomic masses are used, but the quantities and units are different. For example, the relative molecular mass of water is 18 (dimensionless), while the molar mass of water is approximately 18 g mol^{-1}.

Molar mass is the conversion factor between mass in grams and amount in moles. To find moles, divide mass by molar mass. To find mass, multiply amount by molar mass.

Quantity to find	Relationship	Rearrangement
Amount, n	n = m ÷ M	mass divided by molar mass
Mass, m	m = n × M	moles multiplied by molar mass
Molar mass, M	M = m ÷ n	mass divided by amount

1.1.5 Converting Mass and Moles

A calculation should begin by identifying the quantity given and the quantity required. Write the formula, substitute values with units, calculate, and check whether the answer is reasonable. If the mass is smaller than the molar mass, the amount should be less than one mole; if the mass is twice the molar mass, the amount is two moles.

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Worked Example 4: Mass to moles

Question: Calculate the amount of magnesium in 12.0 g of magnesium. Use M(Mg) = 24.0 g mol^{-1}.

Formula: n = m ÷ M. Substitution: n = 12.0 g ÷ 24.0 g mol^{-1} = 0.500 mol Mg.

Worked Example 5: Moles to mass

Question: Find the mass of 0.250 mol of calcium carbonate, $CaCO_3$. Use Ca = 40, C = 12 and O = 16.

M($CaCO_3$) = 40 + 12 + (3 × 16) = 100 g mol^{-1}. Mass = n × M = 0.250 mol × 100 g mol^{-1} = 25.0 g.

Worked Example 6: Amount of a compound

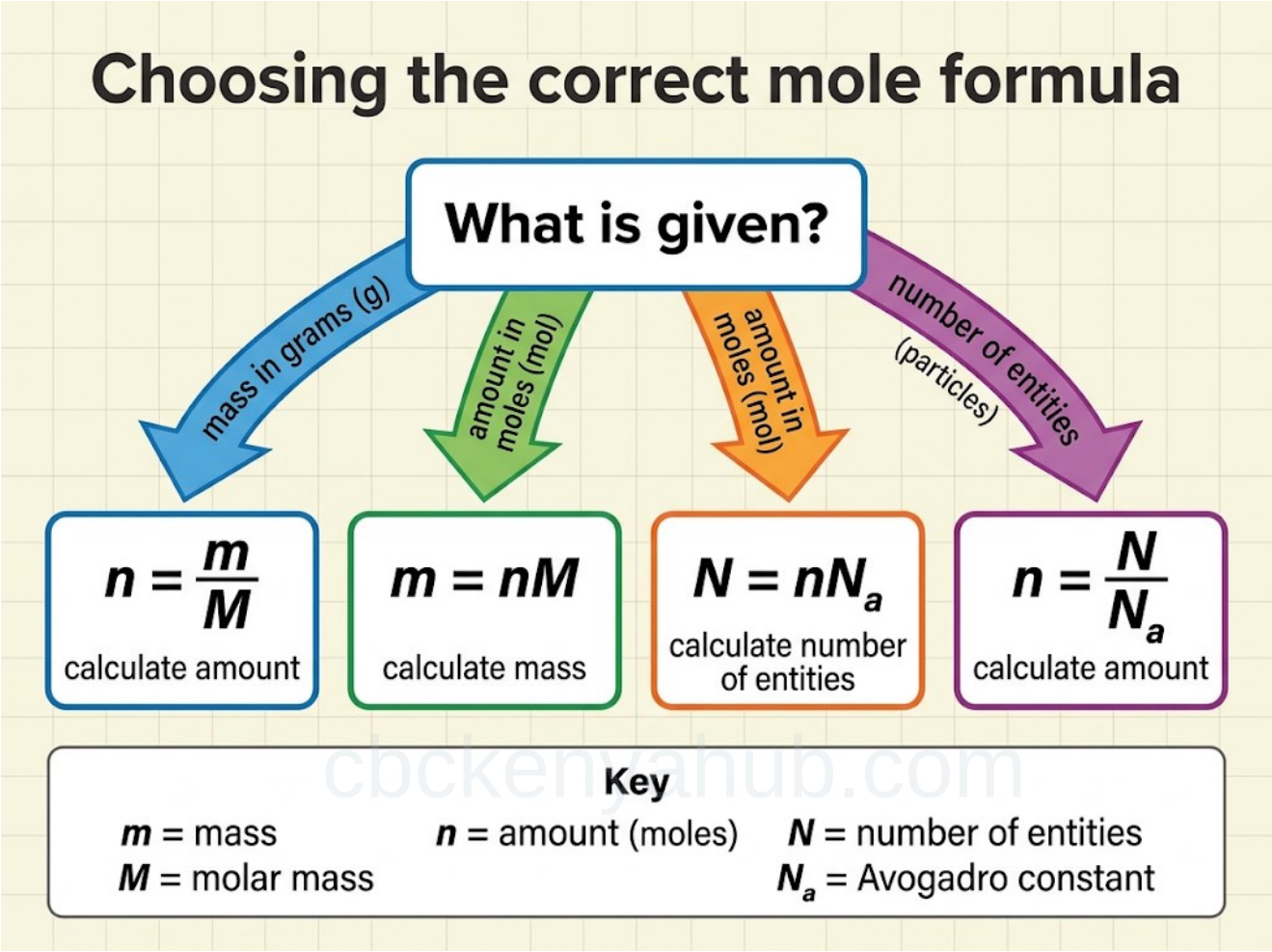
Question: How many moles are present in 9.80 g of sulfuric acid, H_2SO_4? Use H = 1, S = 32 and O = 16.

M(H_2SO_4) = (2 × 1) + 32 + (4 × 16) = 98 g mol^{-1}. n = 9.80 ÷ 98 = 0.100 mol H_2SO_4.

Common errors to avoid

- Using relative formula mass as if it were already in grams; calculate molar mass in g mol^{-1} for mass–mole conversions.
- Multiplying instead of dividing when converting mass to moles.
- Ignoring brackets or subscripts when calculating formula mass.

- Leaving out units or failing to name the entity in a particle calculation.
- Rounding intermediate values too early. Keep extra digits during working and round the final answer sensibly.



1.1.6 Moles and Number of Particles

Once the amount in moles is known, the number of specified entities can be calculated using $N = nN_a$. The reverse calculation is $n = N/N_a$. In a compound, the chemical formula can then be used to determine the number of atoms of each element. For example, each molecule of CO_2 contains one carbon atom and two oxygen atoms; therefore, a sample containing a given number of CO_2 molecules contains twice as many oxygen atoms as carbon atoms.

Worked Example 7: Molecules and atoms

Question: Find the number of molecules in 0.20 mol of carbon dioxide and then find the number of oxygen atoms.

Molecules = $0.20 \times 6.02 \times 10^{23} = 1.204 \times 10^{23}$ molecules. Each molecule has two oxygen atoms, so oxygen atoms = $2 \times 1.204 \times 10^{23} = 2.408 \times 10^{23}$ oxygen atoms.

Worked Example 8: Moles of atoms in a compound

Question: How many moles of hydrogen atoms are present in 0.75 mol of ammonia, NH_3?

Each NH_3 molecule contains three hydrogen atoms. Therefore, 1 mol NH_3 contains 3 mol of hydrogen atoms, and 0.75 mol NH_3 contains $0.75 \times 3 = 2.25$ mol hydrogen atoms.

This relationship is a mole ratio taken directly from the chemical formula. It is different from the ratio of masses because hydrogen and nitrogen atoms have different masses.

1.1.7 Empirical Formula and Molecular Formula: A Useful Extension

The curriculum's core focus is the mole and its use in quantitative chemistry. A useful application of mole calculations is finding the simplest ratio of atoms in a compound. An empirical formula shows the simplest whole-number ratio of atoms. A molecular formula shows the actual number of each type of atom in one molecule.

To obtain an empirical formula from masses or percentages: convert each element's mass to moles; divide each mole value by the smallest; and convert the resulting ratio to whole numbers. If a ratio is close to 1.5, multiply all ratios by 2; if close to 1.33, multiply by 3. Use sensible rounding and check that the final ratio fits the data.

Worked Example 9: Empirical formula

Question: A compound contains 24 g of carbon and 4 g of hydrogen. Find its empirical formula. Use C = 12 and H = 1.

Moles of C = 24/12 = 2 mol. Moles of H = 4/1 = 4 mol. Divide by the smaller value: C:H = 2:4 = 1:2. The empirical formula is CH_2.

The empirical formula is not necessarily the molecular formula. A molecular formula may be a whole-number multiple of the empirical formula, depending on the compound's molar mass.

1.1.8 Molar Solutions and Concentration

A solution is a homogeneous mixture consisting of a solute dissolved in a solvent. The solute is the substance being dissolved, while the solvent is the component that dissolves it. In an aqueous solution, water is the solvent. Quantitative solution chemistry requires the amount of solute to be related to the final volume of solution.

Molar concentration, often called molarity, is the amount of dissolved solute in moles per cubic decimetre of solution. Its unit is mol dm^{-3}. The standard relationship is $c = n/V$, where c is concentration in mol dm^{-3}, n is amount in mol and V is volume in dm^3. One dm^3 equals 1 litre, and 1 dm^3 = 1000 cm^3.

Volume	Equivalent volume
1 dm^3	1000 cm^3

0.500 dm^3	500 cm^3
0.250 dm^3	250 cm^3
0.0250 dm^3	25.0 cm^3
1 cm^3	0.001 dm^3

Before using $c = n/V$, ensure that the volume is in dm^3. Divide a volume in cm^3 by 1000 to convert it to dm^3. A concentration of 0.50 mol dm^{-3} means that each dm^3 of solution contains 0.50 mol of solute; it does not mean that 0.50 mol is present in every volume of solution.

Worked Example 10: Concentration from moles

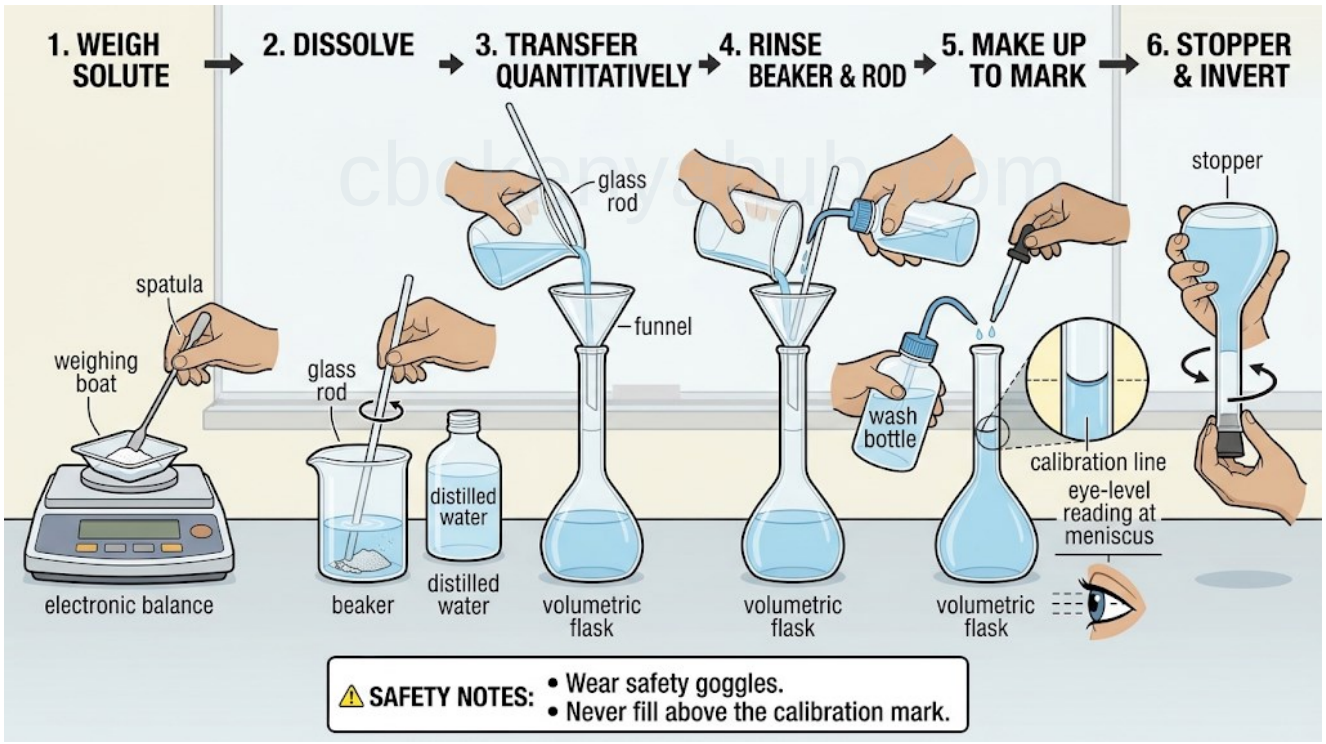
Question: 0.20 mol of sodium chloride is dissolved to make 500 cm^3 of solution. Calculate the concentration.

Convert volume: 500 cm^3 ÷ 1000 = 0.500 dm^3. Then $c = n/V = 0.20/0.500 = 0.40$ mol dm^{-3}.

Worked Example 11: Moles in a measured volume

Question: Calculate the amount of sodium hydroxide in 25.0 cm^3 of 0.100 mol dm^{-3} NaOH.

Volume = 25.0/1000 = 0.0250 dm^3. Amount $n = cV = 0.100 \times 0.0250 = 0.00250$ mol NaOH.



1.1.9 Preparing a Standard Solution

A standard solution has a known concentration. It is prepared carefully so that the amount of solute and final volume are known accurately. The final volume is the volume of the solution, not simply the volume of water initially added.

For a soluble solid, the usual procedure is to calculate the mass required, weigh it accurately, dissolve it in a small volume of distilled water, transfer the solution quantitatively into a volumetric flask, rinse the original container and funnel into the flask, then add distilled water to the calibration mark. The final few drops should be added carefully with a dropper. Stopper the flask and invert it repeatedly to mix the solution.

The word quantitatively means that as much of the weighed solute as possible is transferred to the volumetric flask. If solid or solution is left behind in the beaker or funnel, the actual amount in the flask will be lower than intended and the concentration will be too low.

Worked Example 12: Mass required for a standard solution

Question: Calculate the mass of anhydrous sodium carbonate, Na_2CO_3, needed to prepare 250 cm^3 of 0.100 mol dm^{-3} solution. Use Na = 23, C = 12 and O = 16.

Step 1: Convert volume to dm^3: 250/1000 = 0.250 dm^3. Step 2: Calculate amount: $n = cV = 0.100 \times 0.250 = 0.0250$ mol. Step 3: Calculate molar mass: $M(Na_2CO_3) = (2 \times 23) + 12 + (3 \times 16) = 106$ g mol^{-1}. Step 4: Mass = $nM = 0.0250 \times 106 = 2.65$ g.

Therefore, weigh 2.65 g of anhydrous sodium carbonate, dissolve it, transfer quantitatively and make the solution up to 250 cm^3 in a volumetric flask.

Making a solution by dilution

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Dilution lowers concentration by adding solvent without changing the amount of solute transferred. Since the amount of solute before and after dilution is the same, $C_1V_1 = C_2V_2$. Concentrations may use the same units and volumes may use the same units on both sides; the key is consistency.

Worked Example 13: Dilution

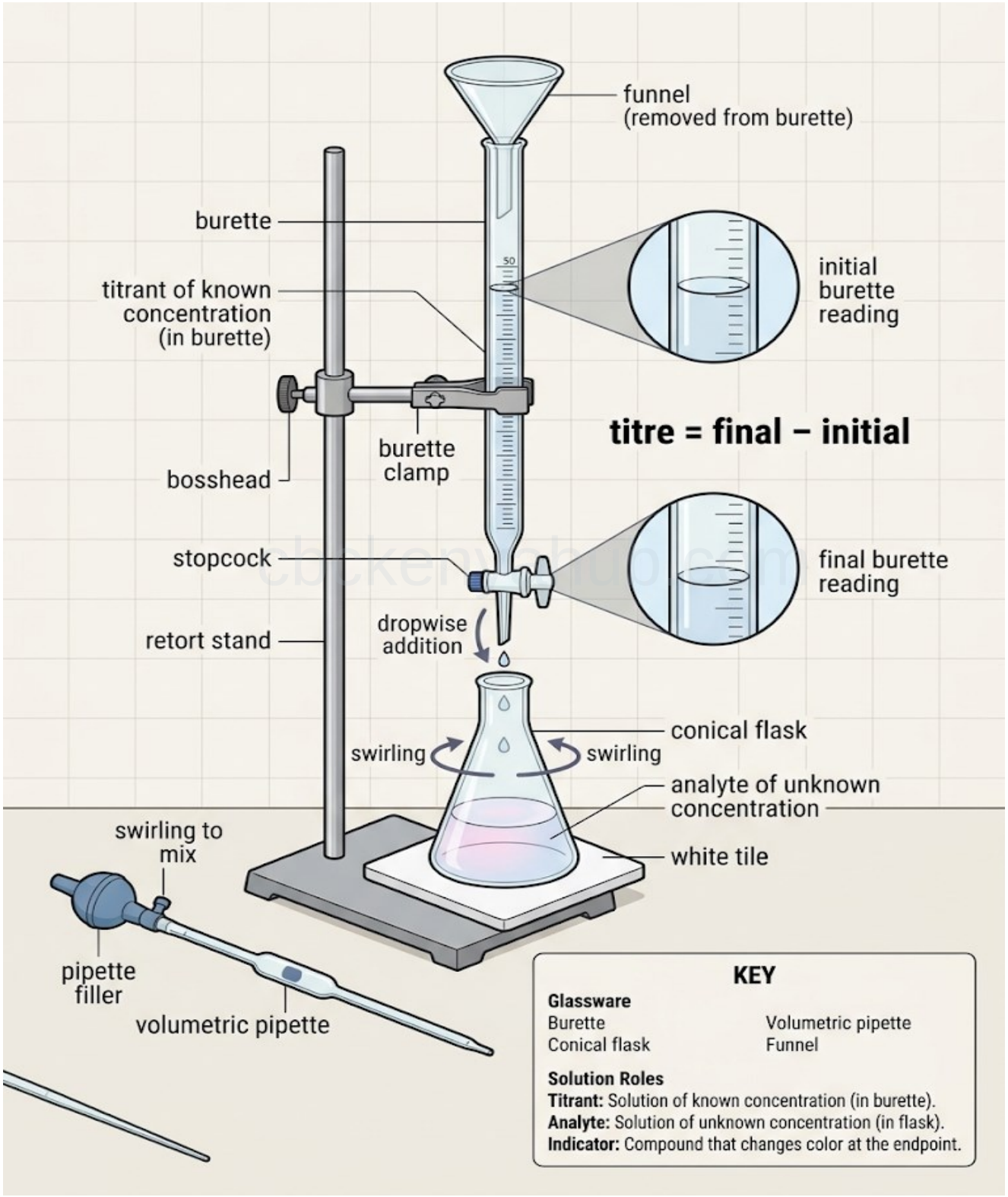
Question: What volume of 2.00 mol dm^{-3} stock solution is required to prepare 250 cm^3 of 0.200 mol dm^{-3} solution?

$V_1 = C_2V_2/C_1 = (0.200 \times 250)/2.00 = 25.0$ cm^3. Measure 25.0 cm^3 of stock solution using suitable volumetric apparatus and dilute to a final volume of 250 cm^3.

1.1.10 Titration: Purpose and Principle

Titration is a quantitative technique used to determine the concentration of a solution by reacting a measured volume of it with a solution of known concentration. In an acid–base titration, an acid reacts with a base according to a balanced chemical equation. The endpoint is detected using a suitable indicator or an instrumental method; the equivalence point is the point at which reactants have combined in the exact stoichiometric ratio shown by the equation.

The burette delivers a measured variable volume of titrant. The pipette transfers a fixed accurate volume of the analyte into a conical flask. A suitable indicator is added when appropriate, and the titrant is delivered while the flask is swirled. Near the endpoint, titrant is added dropwise. A concordant titre is one that agrees closely with other reliable titres, according to the accepted tolerance used in the school laboratory.



Essential titration terms

- Titrant: the solution delivered from the burette, usually of known concentration.
- Analyte: the solution being analysed in the conical flask.
- Aliquot: the accurately measured portion transferred by pipette.
- Titre: the volume delivered from the burette, calculated as final reading minus initial reading.
- Indicator: a substance that changes colour over a characteristic pH range.
- Concordant titres: closely agreeing titre values used to obtain a reliable mean.

The burette reading is usually recorded to two decimal places when the apparatus supports that precision. Read the bottom of the meniscus for most colourless aqueous solutions, with the eye level with the meniscus to avoid parallax error. Some coloured solutions require a different reading convention according to the apparatus and laboratory instructions.

1.1.11 Acid–Base Stoichiometry

The balanced equation determines the mole ratio between acid and base. For a monoprotic acid reacting with a monobasic hydroxide, the ratio may be 1:1, but this should never be assumed for every acid–base reaction. Sulfuric acid can provide two acidic hydrogen ions per molecule in full neutralisation, and calcium hydroxide contains two hydroxide ions per formula unit.

A dependable titration calculation follows this order: write the balanced equation; calculate the moles of the solution with known concentration; use the coefficient ratio to find moles of the unknown; divide by the volume of the unknown in dm^3 to find its concentration.

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Worked Example 14: 1:1 neutralisation

Question: 25.0 cm^3 of hydrochloric acid is neutralised by 20.0 cm^3 of 0.100 mol dm^{-3} sodium hydroxide. Calculate the acid concentration.

Equation: $HCl + NaOH \rightarrow NaCl + H_2O$. Moles NaOH = $0.100 \times 0.0200 = 0.00200$ mol. The ratio HCl:NaOH is 1:1, so moles HCl = 0.00200 mol. Volume HCl = 0.0250 dm^3. Concentration HCl = 0.00200/0.0250 = 0.0800 mol dm^{-3}.

1.1.12 Acid–Base Titration: Worked Examples

Worked Example 15: Sulfuric acid and sodium hydroxide

Question: 25.0 cm^3 of sulfuric acid is completely neutralised by 30.0 cm^3 of 0.200 mol dm^{-3} sodium hydroxide. Calculate the concentration of the sulfuric acid.

Balanced equation: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$.

Moles NaOH = $0.200 \times 0.0300 = 0.00600$ mol. The equation shows that 2 mol NaOH react with 1 mol H_2SO_4, so moles $H_2SO_4 = 0.00600 \div 2 = 0.00300$ mol. Volume acid = 25.0 cm^3 = 0.0250 dm^3. Concentration H_2SO_4 = 0.00300/0.0250 = 0.120 mol dm^{-3}.

Worked Example 16: Calcium hydroxide and nitric acid

Question: 20.0 cm^3 of calcium hydroxide solution reacts completely with 25.0 cm^3 of 0.0800 mol dm^{-3} nitric acid. Find the concentration of $Ca(OH)_2$.

Equation: $Ca(OH)_2 + 2HNO_3 \rightarrow Ca(NO_3)_2 + 2H_2O$. Moles $HNO_3 = 0.0800 \times 0.0250 = 0.00200$ mol. Moles $Ca(OH)_2 = 0.00200 \div 2 = 0.00100$ mol. Volume $Ca(OH)_2$ = 0.0200 dm^3. Concentration = 0.00100/0.0200 = 0.0500 mol dm^{-3}.

Choosing an indicator

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The indicator should change colour over a pH range that gives a clear endpoint close to the equivalence point. Phenolphthalein is commonly used for strong acid–strong base titrations and for many titrations where the endpoint is on the alkaline side; methyl orange is useful where the endpoint is on the acidic side. The correct choice depends on the titration curve and the reagents. In a strong acid–strong base titration, several common indicators can work because the pH changes sharply near equivalence.

Indicator	Acidic solution	Alkaline solution	Typical use note
Methyl orange	Red	Yellow	Useful for endpoints in the acidic transition range
Phenolphthalein	Colourless	Pink	Useful for endpoints in the alkaline transition range
Litmus	Red	Blue	Broad test for acidity/alkalinity; less precise for quantitative titration

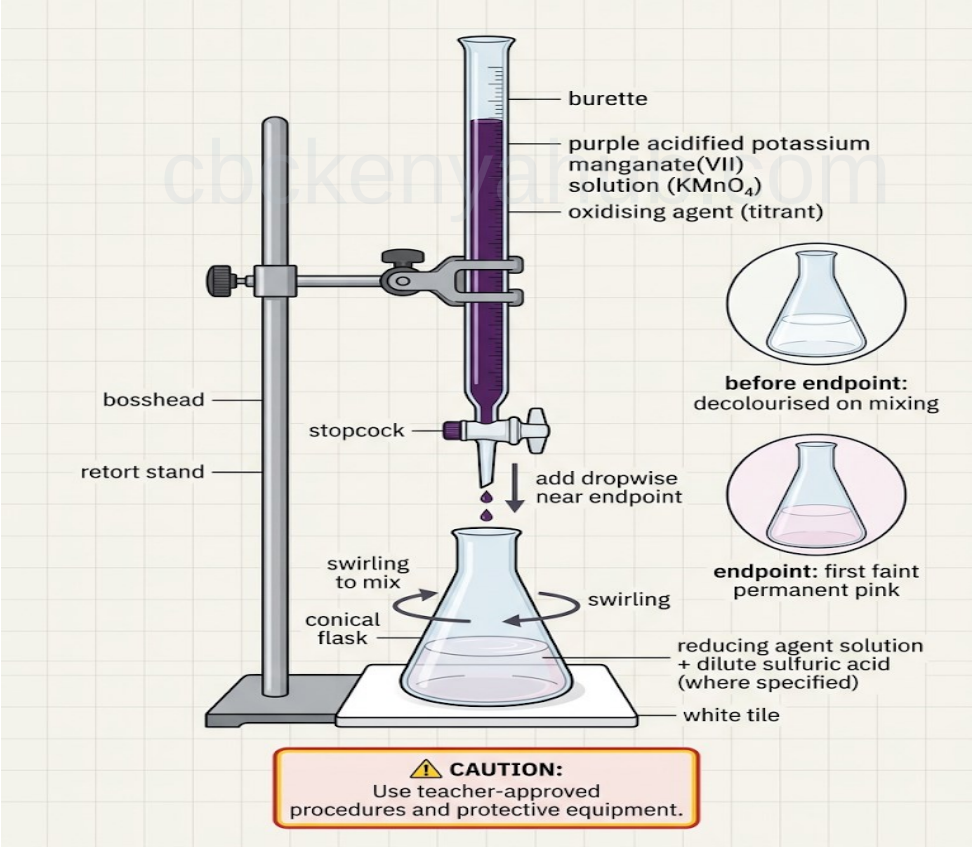
A good technique improves accuracy: rinse the burette with titrant and the pipette with the solution to be measured; rinse the conical flask with distilled water only; remove the funnel from the burette before taking readings; and record initial and final readings. Rinsing the conical flask with distilled water does not change the amount of analyte delivered by the pipette.

1.1.13 Redox Titrations

Oxidation–reduction (redox) reactions involve electron transfer. Oxidation is loss of electrons, while reduction is gain of electrons. In a redox titration, a solution of known concentration reacts with an analyte through a redox process. The balanced redox equation gives the mole ratio needed for the calculation.

Potassium manganate(VII), $KMnO_4$, is a strong oxidising agent often used in acidic solution. Its purple colour can act as its own indicator: the endpoint is commonly the first faint pink colour that persists after swirling. The reaction conditions matter. In acidic solution, manganate(VII) ions are generally reduced to Mn^{2+}; in neutral or alkaline conditions, different products may form. Sulfuric acid is commonly used to acidify the solution in suitable school experiments because hydrochloric acid may be oxidised under some conditions and nitric acid is itself an oxidising agent.

Another common oxidising agent is potassium dichromate(VI), $K_2Cr_2O_7$, which is orange and is reduced to green Cr^{3+} in acidic solution. It generally requires a suitable indicator because its colour change is less convenient for self-indication. Chromium(VI) compounds are hazardous and should only be handled under strict laboratory controls.



Worked Example 17: Permanganate and iron(II)

In acidic solution, the ionic equation is: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$.

Question: 25.0 cm^3 of iron(II) solution requires 20.0 cm^3 of 0.0200 mol dm^{-3} $KMnO_4$. Find the concentration of Fe^{2+}.

Moles MnO_4^- = 0.0200 × 0.0200 = 0.000400 mol. The ratio MnO_4^-:Fe^{2+} is 1:5, so moles Fe^{2+} = 5 × 0.000400 = 0.00200 mol. Volume Fe^{2+} = 0.0250 dm^3. Concentration Fe^{2+} = 0.00200/0.0250 = 0.0800 mol dm^{-3}.

1.1.14 Mole Ratios in Chemical Equations

A balanced chemical equation expresses conservation of atoms and gives the stoichiometric ratio in which reactants combine and products form. The coefficients refer to relative numbers of particles or relative amounts in moles; they do not directly give mass ratios unless the molar masses are also taken into account.

For example, $N_2 + 3H_2 \rightarrow 2NH_3$ means that 1 mol of nitrogen reacts with 3 mol of hydrogen to produce 2 mol of ammonia, assuming the reaction proceeds according to the equation. The same ratio applies to molecules, but gas volumes have this ratio only when gases are measured at the same temperature and pressure.

Worked Example 18: Reactant-to-product calculation

Question: What amount of ammonia can theoretically form from 4.0 mol of hydrogen if nitrogen is in excess?

Equation: $N_2 + 3H_2 \rightarrow 2NH_3$. The ratio H_2:NH_3 is 3:2. Amount NH_3 = 4.0 × 2/3 = 2.67 mol NH_3.

Worked Example 19: Limiting reactant

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Question: 2.0 mol of nitrogen reacts with 5.0 mol of hydrogen. Identify the limiting reactant and calculate the theoretical amount of ammonia.

Equation: $N_2 + 3H_2 \rightarrow 2NH_3$. Two moles of N_2 would require 6.0 mol H_2, but only 5.0 mol H_2 is available. Hydrogen is limiting. From the ratio 3 mol H_2 : 2 mol NH_3, theoretical NH_3 = 5.0 × 2/3 = 3.33 mol.

The limiting reactant is consumed first and limits the maximum possible amount of product. The reactant left over is in excess. In practical reactions, the actual yield may be lower than the theoretical yield because reactions may be incomplete, side reactions may occur, or product may be lost during separation.

Worked Example 20: Mass of product

Question: Calculate the mass of magnesium oxide formed when 4.8 g of magnesium burns completely in oxygen. Use Mg = 24 and O = 16.

Equation: $2Mg + O_2 \rightarrow 2MgO$. Moles Mg = 4.8/24 = 0.20 mol. The ratio Mg:MgO is 1:1, so moles MgO = 0.20 mol. M(MgO) = 24 + 16 = 40 g mol^{-1}. Mass MgO = 0.20 × 40 = 8.0 g.

1.1.15 Percentage Yield and Purity

Theoretical yield is the maximum quantity of product predicted by stoichiometric calculations from the limiting reactant. Actual yield is the amount obtained experimentally. Percentage yield compares the actual yield with the theoretical yield.

Percentage yield = (actual yield ÷ theoretical yield) × 100%. A yield below 100% may result from incomplete reaction, competing reactions, product loss during filtration or transfer, or measurement uncertainty. An apparent yield above 100% often indicates that the product is wet, contaminated, or measured inaccurately.

Worked Example 21: Percentage yield

Question: A reaction has a theoretical yield of 12.0 g but produces 9.0 g of purified product. Calculate the percentage yield.

Percentage yield = (9.0 ÷ 12.0) × 100% = 75%.

Purity and mass calculations

A sample may contain the desired substance mixed with impurities. If the percentage purity is known, the mass of pure substance is calculated as: mass of pure substance = (percentage purity ÷ 100) × mass of impure sample. Only the pure reactive component should be used in stoichiometric calculations.

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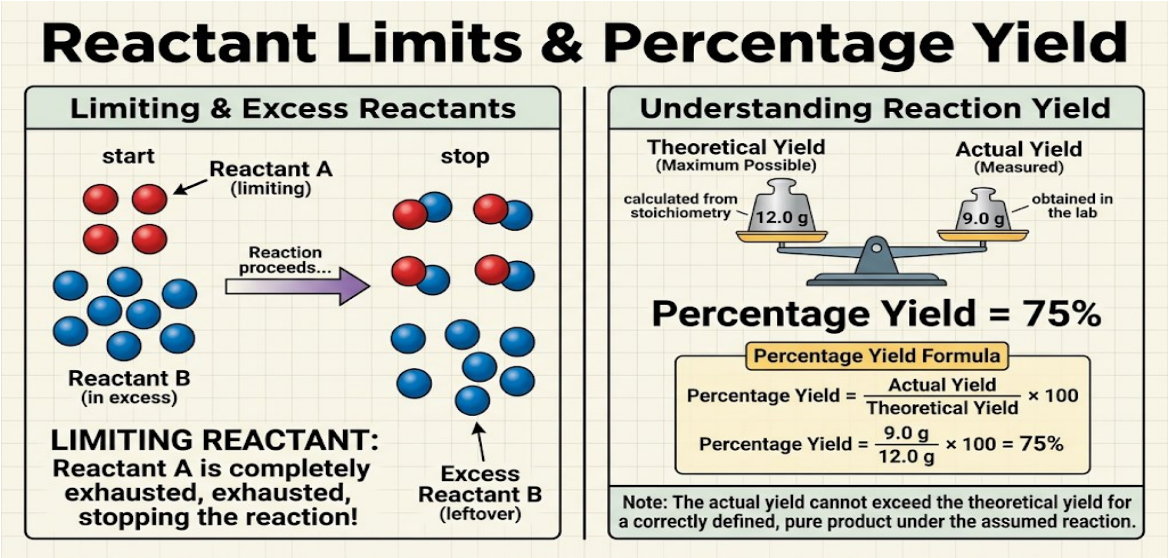
Worked Example 22: Impure reactant

Question: A 10.0 g sample of limestone is 80.0% calcium carbonate. Find the mass of $CaCO_3$ in the sample.

Mass of $CaCO_3$ = 0.800 × 10.0 g = 8.00 g. The remaining 2.00 g consists of impurities. In any reaction calculation, use 8.00 g as the mass of calcium carbonate, not 10.0 g.



1.1.16 Molar Gas Volume

Gas volume depends on temperature and pressure. At the same temperature and pressure, equal volumes of gases contain equal numbers of molecules, provided the gases behave approximately ideally. This makes it possible to compare gaseous reactants and products using volume ratios when the conditions are the same.

Molar gas volume is the volume occupied by one mole of gas at a specified temperature and pressure. There is no single molar volume that applies at all conditions. A commonly used approximate value for gases near room temperature and pressure is about 24 $dm^3\ mol^{-1}$, while at 0 °C and 1 atmosphere the ideal-gas molar volume is approximately 22.4 $dm^3\ mol^{-1}$. Follow the value and conditions stated in the question or prescribed by the teacher.

If the molar gas volume V_m is given, amount can be calculated using $n = V/V_m$ and volume using $V = nV_m$. Ensure that the units match: for example, use volume in dm^3 with V_m in $dm^3\ mol^{-1}$.

Worked Example 23: Volume from moles

Question: Calculate the volume occupied by 0.50 mol of oxygen at conditions where the molar gas volume is 24.0 $dm^3\ mol^{-1}$.

$V = nV_m = 0.50 \times 24.0 = 12.0\ dm^3$.

Worked Example 24: Moles from gas volume

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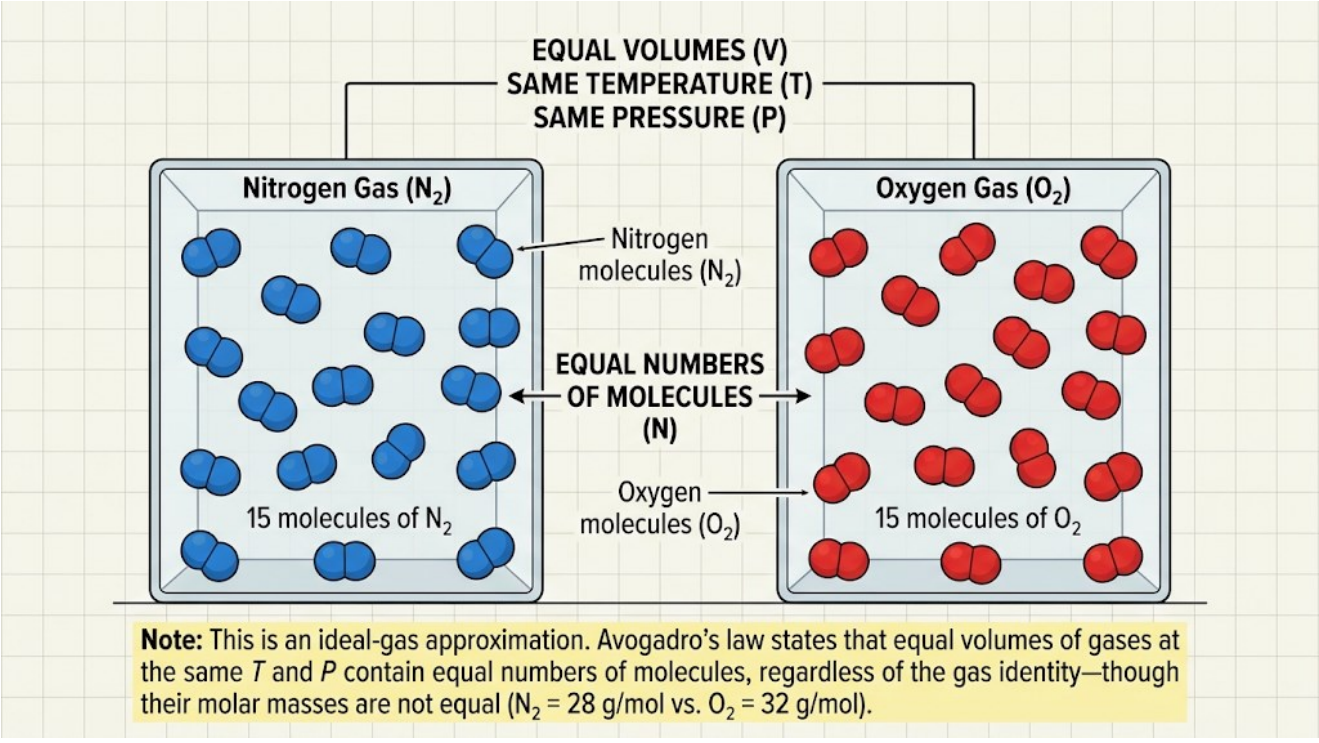
Question: Calculate the amount of carbon dioxide in 600 cm^3 of gas at conditions where $V_m = 24.0\ dm^3\ mol^{-1}$.

Convert 600 cm^3 to 0.600 dm^3. Then $n = V/V_m = 0.600/24.0 = 0.0250$ mol CO_2.



1.1.17 Gas Stoichiometry

When gaseous reactants and products are measured at the same temperature and pressure, their volume ratios follow the coefficients in the balanced equation. This follows from the relationship between gas volume and amount under the same conditions. The ratio is not a mass ratio.

Worked Example 25: Hydrogen and oxygen

Question: What volume of oxygen reacts completely with 60 cm^3 of hydrogen, and what volume of steam forms, if all gases are compared at the same temperature and pressure?

Equation: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$. The volume ratio is 2:1:2. Oxygen volume = 60 ÷ 2 = 30 cm^3. Steam volume = 60 cm^3, assuming the water remains gaseous at the stated conditions.

Worked Example 26: Ammonia formation

Question: Nitrogen and hydrogen react according to $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. Calculate the volume of hydrogen required to react with 40 cm^3 of nitrogen and the volume of ammonia produced, assuming complete reaction at the same temperature and pressure.

The ratio N_2:H_2:NH_3 is 1:3:2. Hydrogen required = 3 × 40 = 120 cm^3. Ammonia produced = 2 × 40 = 80 cm^3.

1.1.18 Using the Ideal Gas Equation

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Where a question supplies pressure, volume, temperature and amount, the ideal gas equation may be used: $pV = nRT$. Here p is pressure, V is volume, n is amount, T is absolute temperature in kelvin, and R is the gas constant. Units must be consistent with the selected value of R. Temperature in kelvin is found by adding approximately 273 to the Celsius temperature.

At Grade 11, the key practical lesson is that gas volume changes when temperature or pressure changes. A calculation should use the exact conditions and constants provided in the question. Do not mix the molar volume at one set of conditions with measurements taken at another.

Worked Example 27: Temperature conversion

Question: Convert 27 °C to kelvin.

$T(K) = t(°C) + 273.15 = 27 + 273.15 = 300.15$ K, usually reported as approximately 300 K for a whole-degree starting value.

1.1.19 Problem-Solving Framework for Mole Questions

Many chemistry calculations become straightforward when the same sequence is used consistently. First, identify what is given and what is required. Second, write the relevant equation and balance any chemical reaction. Third, convert the given measurement to moles if necessary. Fourth, use the balanced equation or formula to apply the correct mole ratio. Fifth, convert to the required unit and check the result.

Question type	Typical route	Check
Mass → amount	$n = m/M$	Result in mol
Amount → mass	$m = nM$	Result in g
Particles → amount	$n = N/N_a$	Entity specified
Amount → particles	$N = nN_a$	Atoms/molecules/ions stated
Solution concentration	$c = n/V$	V in dm^3
Titration	cV → mole ratio → unknown c	Balanced equation used
Gas volume	$n = V/V_m$ or $V = nV_m$	Same stated conditions
Reaction mass	mass → moles → mole ratio → moles → mass	Limiting reactant checked

Mixed Worked Example 28: Multi-step reaction

Question: 5.6 g of iron reacts completely with sulfur to form iron(II) sulfide. Calculate the mass of iron(II) sulfide formed. Use Fe = 56, S = 32.

Equation: Fe + S → FeS. Moles Fe = 5.6/56 = 0.10 mol. Ratio Fe:FeS = 1:1, so moles FeS = 0.10 mol. M(FeS) = 56 + 32 = 88 g mol^{-1}. Mass FeS = 0.10 × 88 = 8.8 g.

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Mixed Worked Example 29: Concentration and dilution

Question: 10.0 cm^3 of 1.50 mol dm^{-3} solution is diluted to 100 cm^3. Find the new concentration.

$C_1V_1 = C_2V_2$. $C_2 = C_1V_1/V_2 = (1.50 \times 10.0)/100 = 0.150$ mol dm^{-3}. The concentration becomes one-tenth of the original because the final volume is ten times the initial volume.

1.1.20 Summary: The Mole

- The mole is the SI unit for amount of substance and contains $6.02214076 \times 10^{23}$ specified entities.
- The Avogadro constant is commonly approximated as 6.02×10^{23} mol^{-1}.
- Relative atomic mass, A_r, and relative formula mass, M_r, are ratios and have no units.
- Molar mass is the mass of one mole and is measured in g mol^{-1}.
- Amount = mass ÷ molar mass; mass = amount × molar mass.
- Number of entities = amount × Avogadro constant; amount = number of entities ÷ Avogadro constant.
- Concentration in mol dm^{-3} equals amount in moles divided by volume in dm^3.
- To prepare a standard solution, weigh accurately, dissolve, transfer quantitatively, make up to the mark, and mix thoroughly.

- Dilution follows $C_1V_1 = C_2V_2$ when the same amount of solute is conserved.
- Titration calculations depend on accurate volumes, reliable titres, a suitable endpoint and the balanced equation's mole ratio.
- Redox titrations use electron-transfer reactions; reaction conditions determine the products and colour changes.
- Balanced equations provide mole ratios. Check limiting reactants before calculating theoretical yield.
- Molar gas volume depends on temperature and pressure. Use the value stated for the conditions in the question.
- Always show formula, substitution, units, mole ratio and final answer.

Self-check Questions

1. Calculate the relative formula mass of $Mg(NO_3)_2$ using Mg = 24, N = 14 and O = 16.
2. Find the amount in moles present in 4.60 g of sodium, Na (A_r = 23).
3. How many molecules are present in 0.125 mol of water?
4. Calculate the concentration when 0.0750 mol of solute is dissolved to make 250 cm^3 of solution.
5. What mass of KOH is needed to prepare 500 cm^3 of 0.200 mol dm^{-3} solution? Use K = 39, O = 16, H = 1.
6. 25.0 cm^3 of H_2SO_4 reacts with 40.0 cm^3 of 0.150 mol dm^{-3} NaOH. Calculate the acid concentration.
7. What volume does 0.30 mol of gas occupy when V_m = 24.0 dm^3 mol^{-1}?
8. Explain why a titre is calculated by subtracting the initial burette reading from the final reading.

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Answers for checking: 1) 148; 2) 0.200 mol; 3) 7.53×10^{22} molecules; 4) 0.300 mol dm^{-3}; 5) 5.60 g; 6) 0.120 mol dm^{-3}; 7) 7.20 dm^3; 8) The burette reading measures the delivered volume, which is the change between the final and initial readings.

SUB-STRAND 1.2: NON-METALS

1.2.1 Overview of Non-metals

Non-metals are elements that generally lack the typical metallic properties of good electrical and thermal conductivity, malleability and ductility. Many non-metals are gases at room temperature, while some are solids and one, bromine, is a liquid. Non-metals are found on the right-hand side of the periodic table, with hydrogen as a notable non-metal placed on the left.

The Grade 11 curriculum strand highlights oxygen and nitrogen. Both are important components of air and support essential biological and industrial processes. Oxygen is needed for aerobic respiration and supports combustion, while nitrogen makes up most of the atmosphere and is an essential element in proteins and nucleic acids. Their chemical behaviour differs because their atoms and molecules have different structures and reactivities.

Feature	Oxygen	Nitrogen
Formula of common elemental gas	O_2	N_2
Approximate proportion in dry air	About 21% by volume	About 78% by volume
Appearance at room temperature	Colourless gas	Colourless gas
Relative chemical reactivity	Supports combustion; reacts with many elements under suitable conditions	Relatively unreactive at ordinary conditions because N≡N is strong
Important roles	Respiration, oxidation, steelmaking, medical oxygen	Fertilisers, ammonia production, inert atmospheres, food packaging

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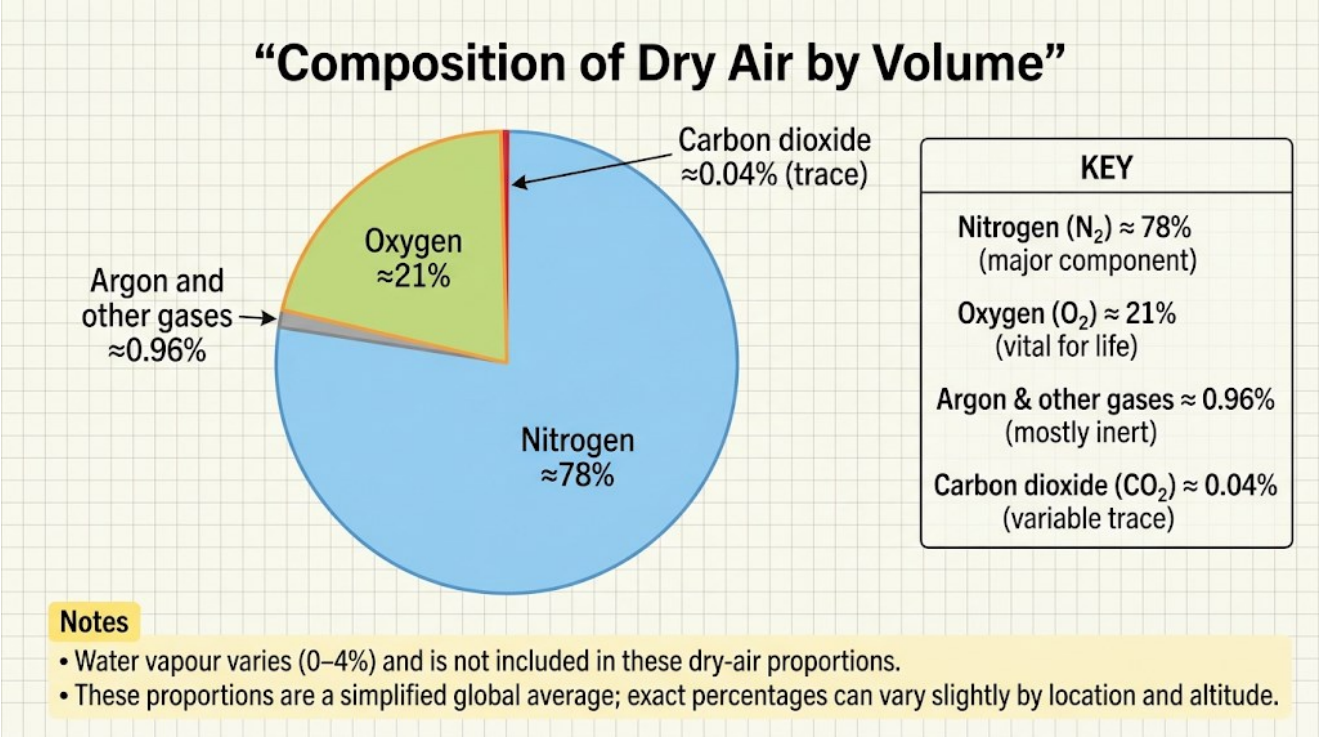


1.2.2 General Properties of Non-metals

Many non-metals have low electrical and thermal conductivity, are brittle when solid, and form acidic oxides. These are generalisations rather than universal rules. Graphite, an allotrope of carbon, conducts electricity; diamond is extremely hard; and some non-metal oxides are neutral rather than acidic. Therefore, classify a substance using observed properties and its chemistry, not a single feature alone.

- Physical state: many are gases at room temperature, but carbon, sulfur, phosphorus and iodine are solids, while bromine is a liquid.
- Appearance: non-metals often lack metallic lustre, although iodine has a shiny surface.
- Conductivity: most are poor conductors, with graphite as a familiar exception.
- Mechanical properties: solid non-metals tend to be brittle rather than malleable or ductile.
- Bonding: non-metals commonly form covalent bonds with other non-metals and can form negative ions in ionic compounds.
- Oxides: many non-metal oxides react with water to form acidic solutions, such as sulfur dioxide forming an acidic solution; exceptions exist.

1.2.3 Oxygen: Occurrence and Importance

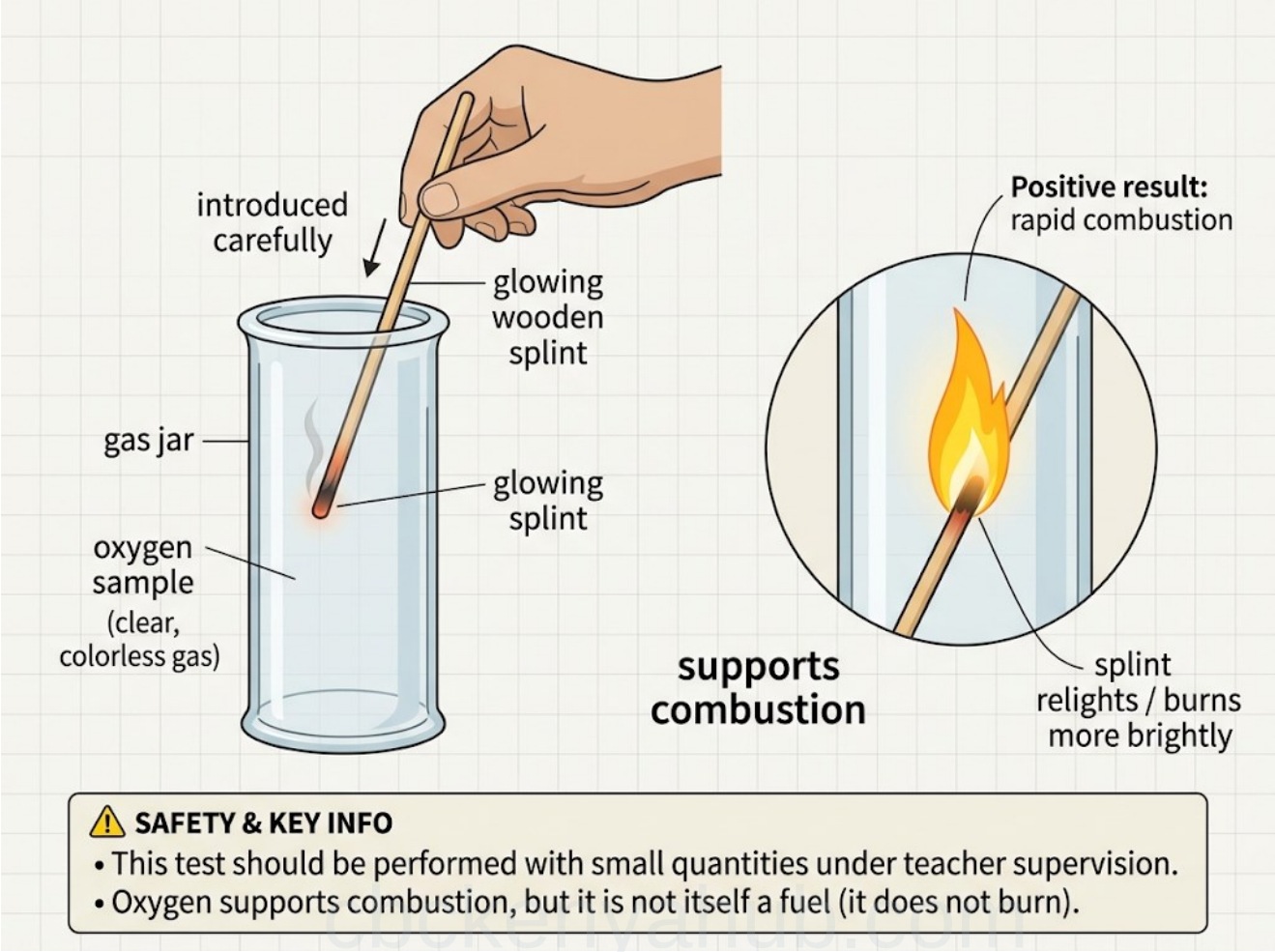
Oxygen occurs naturally in the atmosphere mainly as diatomic oxygen, O_2. It is also chemically combined in water, metal oxides, silicates, carbonates and many organic compounds. Oxygen is abundant in the Earth's crust by mass because it occurs in many minerals and rocks.

Oxygen gas is essential for aerobic respiration, the process by which cells release usable energy from nutrients. In combustion, oxygen acts as an oxidising agent and allows many fuels to burn. Oxygen does not itself burn; rather, it supports the burning of suitable substances. This distinction is important in practical observations.

Oxygen is used in hospitals for patients who require supplemental oxygen, in oxy-fuel cutting and welding, in steel manufacture, and in some chemical processes. Medical oxygen must be supplied using approved equipment and at prescribed flow rates; oxygen-enriched environments greatly increase fire risk.

Physical properties of oxygen

- It is a colourless, odourless gas at room temperature.
- It is only slightly soluble in water, allowing it to be collected over water in appropriate laboratory circumstances.
- It is denser than air under the same temperature and pressure, although collection method should be chosen based on the full apparatus and gas properties.
- It liquefies at very low temperature and is stored as a cryogenic liquid for large-scale industrial distribution.



1.2.4 Chemical Properties of Oxygen

Oxygen reacts with many elements to form oxides. The reaction may be rapid, as in burning magnesium, or slow, as in the rusting of iron. The products depend on the reactant and conditions. In many school experiments, a substance is heated to initiate reaction with oxygen.

Reactant	Example equation	Typical observation / product
Magnesium	$2Mg + O_2 \rightarrow 2MgO$	Bright white flame; white magnesium oxide powder
Carbon	$C + O_2 \rightarrow CO_2$	Carbon dioxide forms when combustion is complete
Sulfur	$S + O_2 \rightarrow SO_2$	Blue flame; sulfur dioxide gas forms
Iron wool	$3Fe + 2O_2 \rightarrow Fe_3O_4$ (simplified common product)	Glowing iron wool; dark iron oxide forms

Oxidation can be described in different ways depending on context. At this level, a common definition is the addition of oxygen to a substance; in redox chemistry, oxidation is more generally loss of electrons. These definitions are consistent in many examples but electron transfer is the broader definition.

Worked explanation: Why does a glowing splint relight?

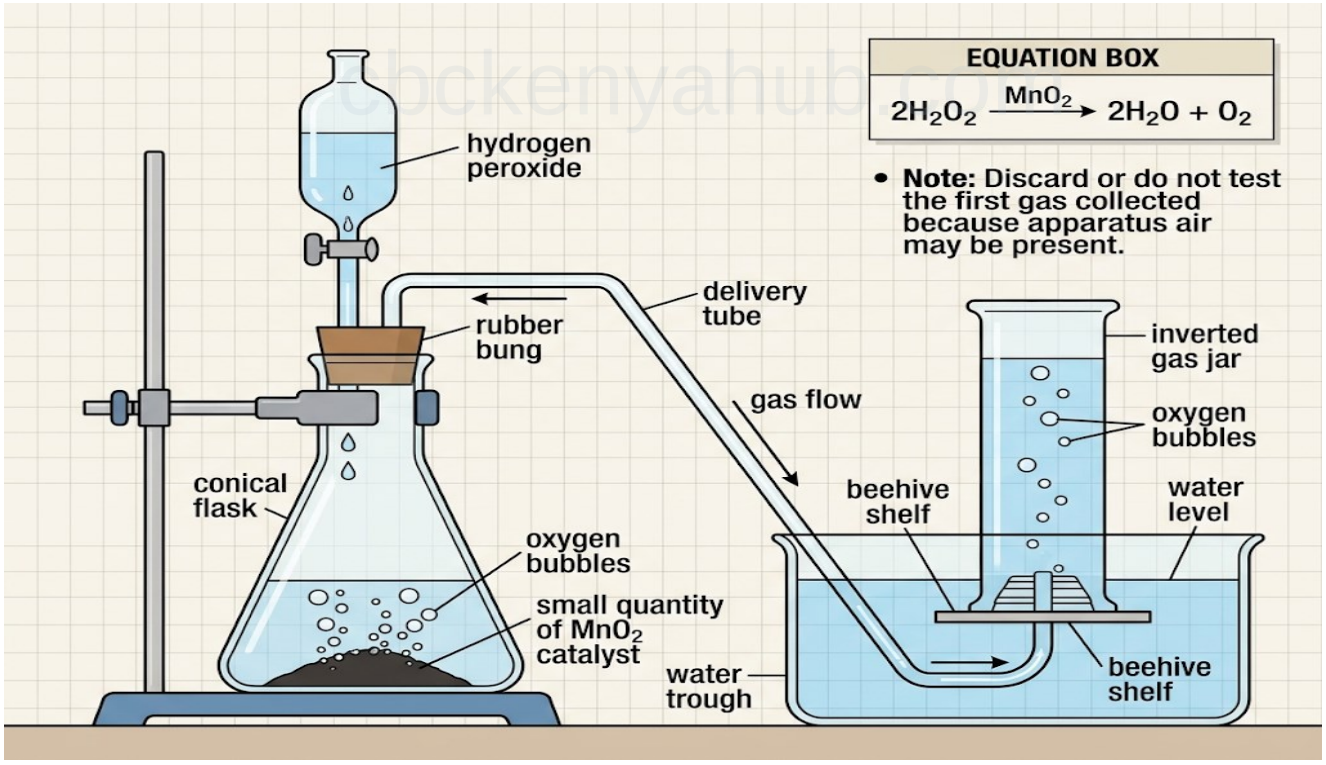
A glowing splint already has a hot, smouldering region. Oxygen increases the rate at which the remaining wood reacts, so the splint may relight or glow more strongly. The result is evidence that the gas supports combustion, but it is not evidence that oxygen is itself flammable.

1.2.5 Laboratory Preparation of Oxygen

In a school laboratory, oxygen may be prepared by decomposing a suitable oxygen-containing compound. One common method uses hydrogen peroxide solution with manganese(IV) oxide as a catalyst. The catalyst increases the reaction rate but is not used up overall.

Equation: $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$. Manganese(IV) oxide, MnO_2, acts as a catalyst. The gas can be collected over water because oxygen is only slightly soluble in water and does not react significantly with it under the collection conditions.

The apparatus usually includes a reaction flask containing hydrogen peroxide, a delivery tube, and a water trough with an inverted gas jar. The first portion of gas is often allowed to escape because it contains the air initially present in the apparatus. The gas jar should be handled carefully and the teacher should specify how much gas to collect.



Alternative method and observations

Oxygen can also be prepared by heating potassium chlorate(V), usually with manganese(IV) oxide as catalyst, under controlled school-laboratory conditions. The equation is $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$. This method involves heating an oxidising solid and requires extra care, appropriate apparatus and teacher supervision.

The main evidence that oxygen has been produced is the relighting of a glowing splint. Gas collection alone does not prove the identity of a gas; confirmatory tests should be performed only using safe, approved procedures.

Advantages of collection over water

- The gas can be collected in a reasonably pure form after the air in the apparatus has been displaced.
- The water level helps show when the gas jar has filled.
- It is appropriate for gases that are only slightly soluble in water and do not react with water.
- It is not appropriate for gases that dissolve readily in water or react with it.

1.2.6 General Methods of Gas Collection

The collection method depends on the gas's solubility in water, density relative to air, and chemical reactivity. A suitable method should minimise contamination and avoid a reaction between the gas and the collection medium.

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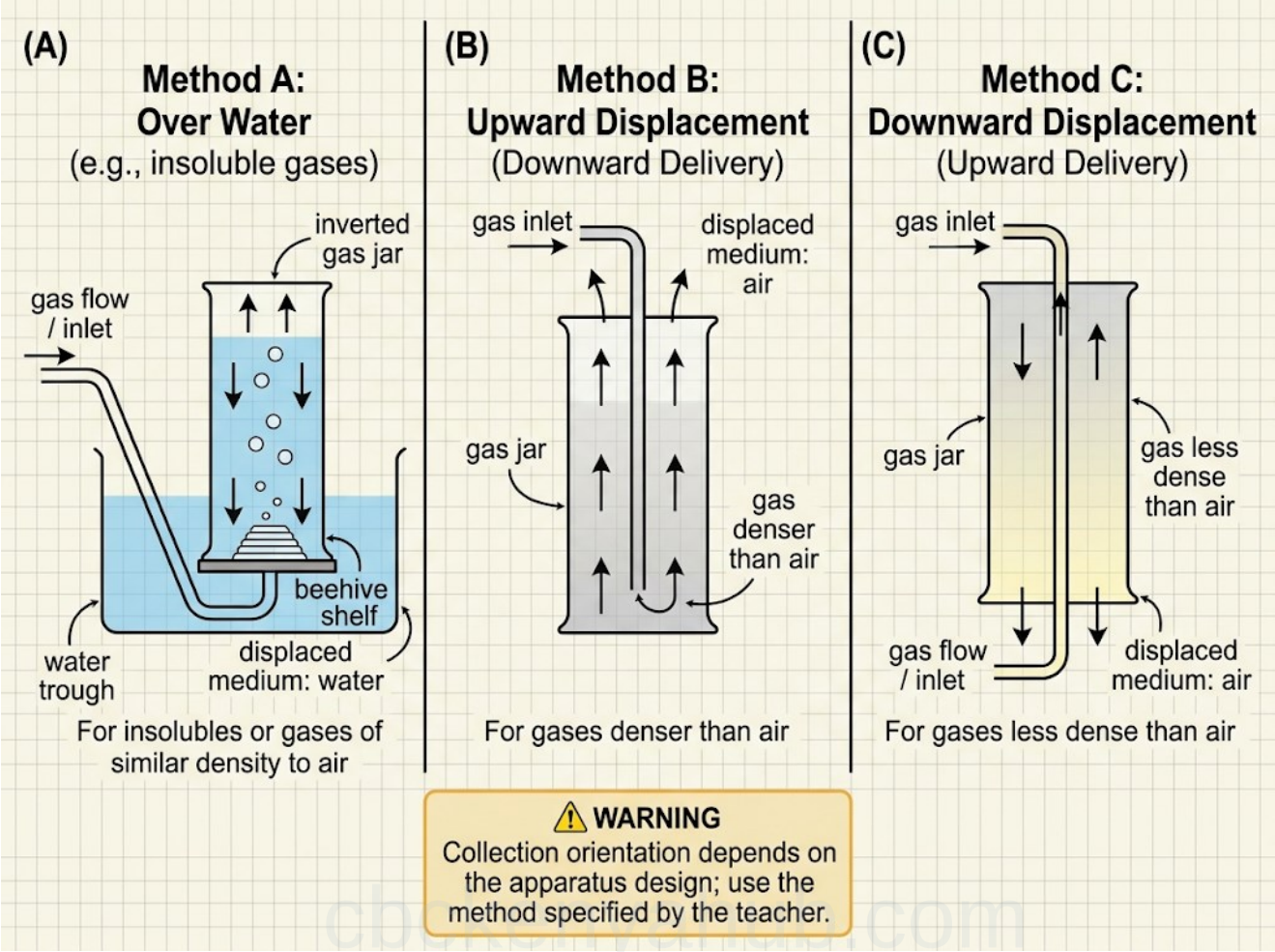
Method	When it is suitable	Important limitation
Collection over water	Gas is only slightly soluble in water and does not react significantly with water	Unsuitable for very soluble gases such as ammonia and hydrogen chloride
Upward delivery / downward displacement of air	Often used for gases denser than air, depending on apparatus arrangement	Air may remain mixed with the sample; avoid using density alone without considering solubility and safety
Downward delivery / upward displacement of air	Often used for gases less dense than air, depending on apparatus arrangement	May give a less pure sample because of mixing with air
Gas syringe	Suitable for measuring gas volume and collecting many gases if materials are compatible	Must be gas-tight and used within its capacity; some gases react with syringe materials

Terminology such as "upward delivery" and "downward delivery" can be confusing because books sometimes describe the direction of gas flow and sometimes the displacement of air. Always interpret the labelled diagram and explain which gas is entering and which fluid or gas is being displaced.



How to choose a method

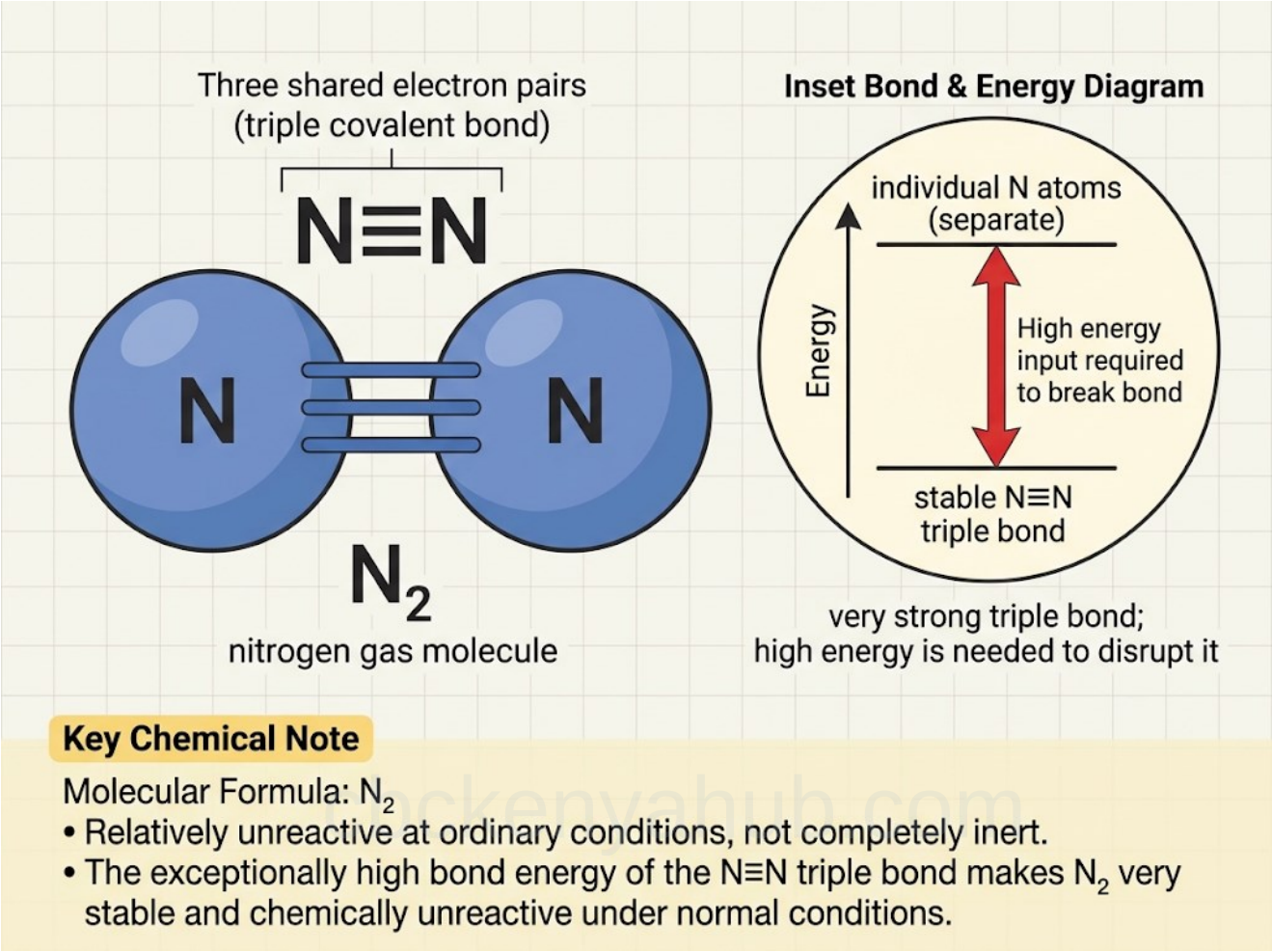
9. Ask whether the gas dissolves in or reacts with water. If it does, collection over water is unsuitable.
10. If water collection is unsuitable, compare the gas density with air and choose a displacement arrangement that minimises mixing.
11. For quantitative volume measurements, use a gas syringe or another calibrated apparatus where appropriate.
12. Consider toxicity, flammability, corrosiveness and pressure. A theoretically suitable method may still be unsafe if the gas is hazardous.

1.2.7 Nitrogen: Occurrence and Properties

Nitrogen is the most abundant gas in dry air, accounting for approximately 78% by volume. It also occurs in nitrates, ammonium salts, proteins and nucleic acids. Although atmospheric nitrogen is abundant, most plants cannot use N_2 gas directly. Nitrogen must be converted into biologically available compounds through processes such as nitrogen fixation, industrial ammonia manufacture, and natural microbial activity.

Nitrogen gas is colourless, odourless and only slightly soluble in water. It is relatively unreactive at ordinary temperatures because the two nitrogen atoms in N_2 are joined by a very strong triple bond. Under suitable

conditions, nitrogen reacts with hydrogen to form ammonia and reacts with oxygen at high temperatures to form nitrogen oxides.



Physical properties of nitrogen

- Colourless and odourless at room temperature.
- Slightly soluble in water.
- Does not support combustion in the way oxygen does.
- Has a density close to, but slightly lower than, dry air under the same conditions.
- Can be liquefied at very low temperatures and separated from other components of air by fractional distillation.

Chemical properties of nitrogen

Nitrogen is not completely unreactive. At high temperatures or under specialised conditions it reacts with oxygen to form nitrogen monoxide: $N_2 + O_2 \rightleftharpoons 2NO$. Nitrogen monoxide can then react with oxygen to form nitrogen dioxide: $2NO + O_2 \rightarrow 2NO_2$. Nitrogen also reacts with hydrogen in the Haber process to form

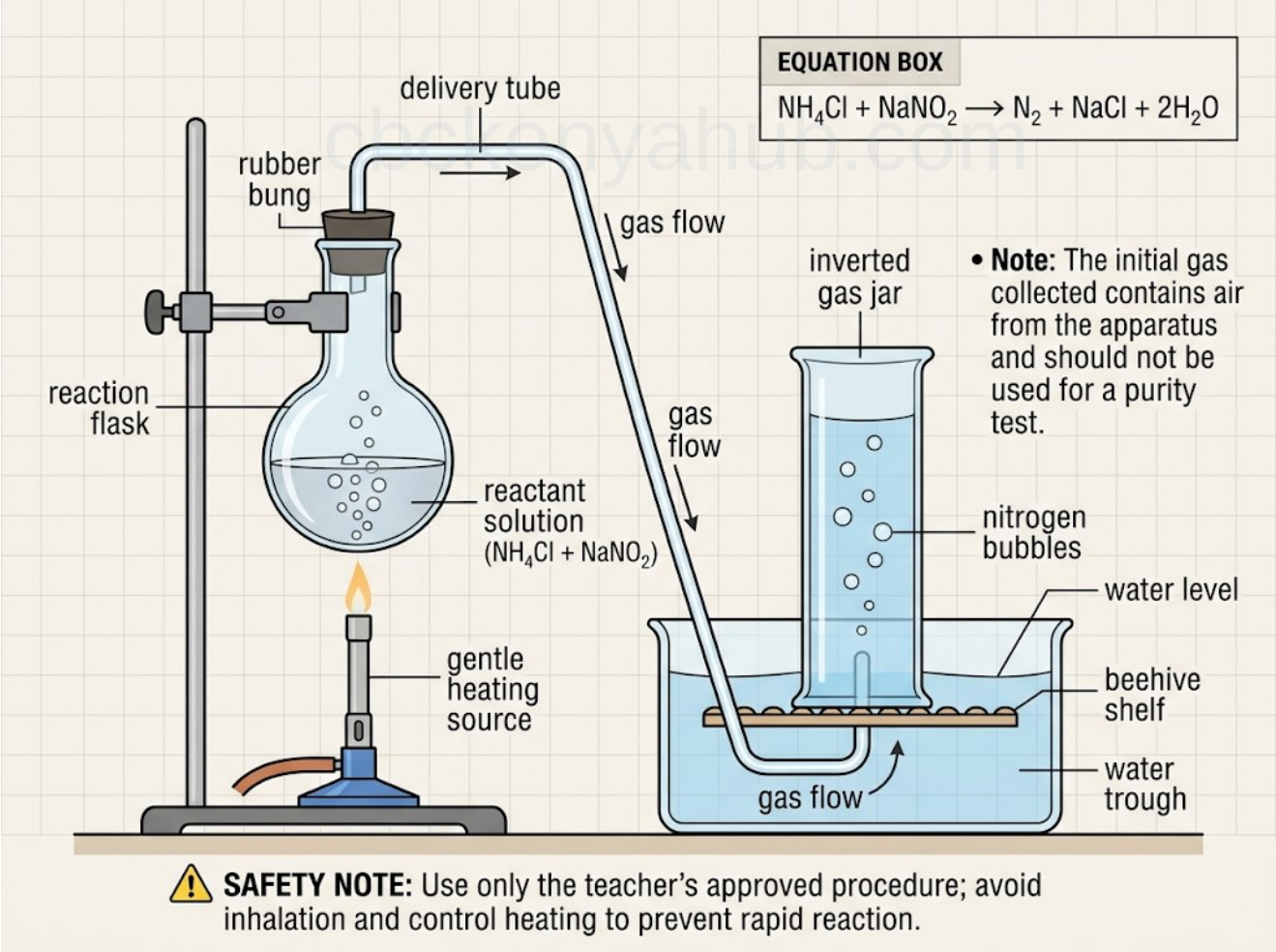
ammonia: $N_2 + 3H_2 \rightleftharpoons 2NH_3$. The Haber process uses elevated temperature and pressure with an iron-based catalyst in an industrial system.

1.2.8 Laboratory Preparation of Nitrogen

Nitrogen may be prepared in the laboratory by removing oxygen and other reactive components from air, or through a suitable chemical reaction selected for the school laboratory. A common textbook approach involves warming a solution containing ammonium chloride and sodium nitrite, which forms ammonium nitrite in situ; ammonium nitrite decomposes to nitrogen and water. This procedure must be carried out only with teacher-approved quantities and apparatus.

Simplified equations: $NH_4Cl + NaNO_2 \rightarrow NH_4NO_2 + NaCl$; then $NH_4NO_2 \rightarrow N_2 + 2H_2O$. The overall reaction is $NH_4Cl + NaNO_2 \rightarrow N_2 + NaCl + 2H_2O$. Actual laboratory instructions may use a different approved preparation method.

Nitrogen can be collected over water because it is only slightly soluble and does not react significantly with water under ordinary collection conditions. The first portion of gas should be treated as contaminated by air in the apparatus.



Testing nitrogen

Nitrogen does not relight a glowing splint and does not support ordinary combustion. However, this test alone cannot positively identify nitrogen because several other gases also fail to support combustion. Reliable identification may require knowing how the gas was prepared and excluding other gases through the experimental procedure.

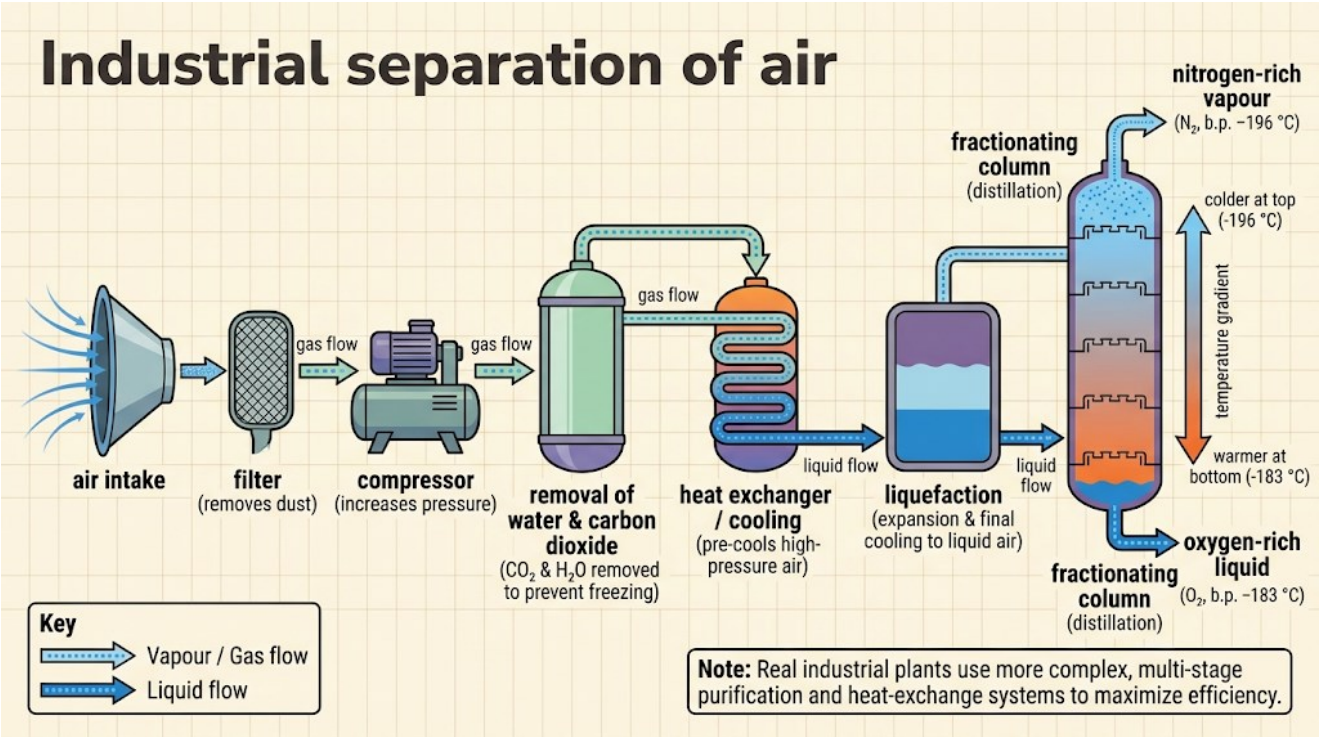
Why the first gas is often discarded

Before gas production begins, the flask and delivery tube contain air. The first gas emerging from the apparatus is therefore a mixture of the prepared gas and the original air. Allowing the initial mixture to escape helps reduce contamination of the collected sample. This principle applies to many gas-preparation experiments.

1.2.9 Industrial Isolation of Oxygen and Nitrogen

Large quantities of oxygen and nitrogen are obtained from air. Air is filtered to remove dust, compressed, cooled and treated to remove water vapour and carbon dioxide, which could freeze and block equipment at low temperatures. The purified air is then cooled until it becomes liquid. Fractional distillation separates the components because they have different boiling points.

Nitrogen has a lower boiling point (about −196 °C) than oxygen (about −183 °C) at atmospheric pressure, so nitrogen is more volatile and boils off earlier in a fractional distillation column. Oxygen-rich liquid remains at lower levels and can be collected separately. Industrial systems use carefully controlled equipment and may include additional purification and separation steps.



Why pre-treatment is necessary

- Dust can damage or obstruct equipment.
- Water can freeze and block pipes at cryogenic temperatures.
- Carbon dioxide can solidify at low temperature and interfere with flow.
- Controlled cooling and heat exchange improve efficiency and support safe operation.

Worked explanation: Which gas distils first?

Nitrogen has the lower boiling point of the two major components. At the same pressure, it changes from liquid to vapour at a lower temperature than oxygen. In the fractional distillation column, nitrogen is therefore more volatile and is preferentially collected toward the top, while oxygen-rich liquid is collected lower down. The explanation is based on boiling point, not on the percentage of each gas in air.

1.2.10 Uses of Oxygen and Nitrogen

Uses of oxygen

- Medical oxygen: supplied to patients who need supplemental oxygen, using prescribed equipment and controlled delivery.
- Metal fabrication: oxygen supports high-temperature oxy-fuel cutting and welding when used with a suitable fuel gas.
- Steelmaking: oxygen is used in controlled oxidation processes to remove unwanted impurities from molten iron.
- Chemical manufacture: oxygen is used as an oxidant in several industrial processes.
- Wastewater treatment and aquaculture: oxygen may be supplied to increase dissolved oxygen when needed.

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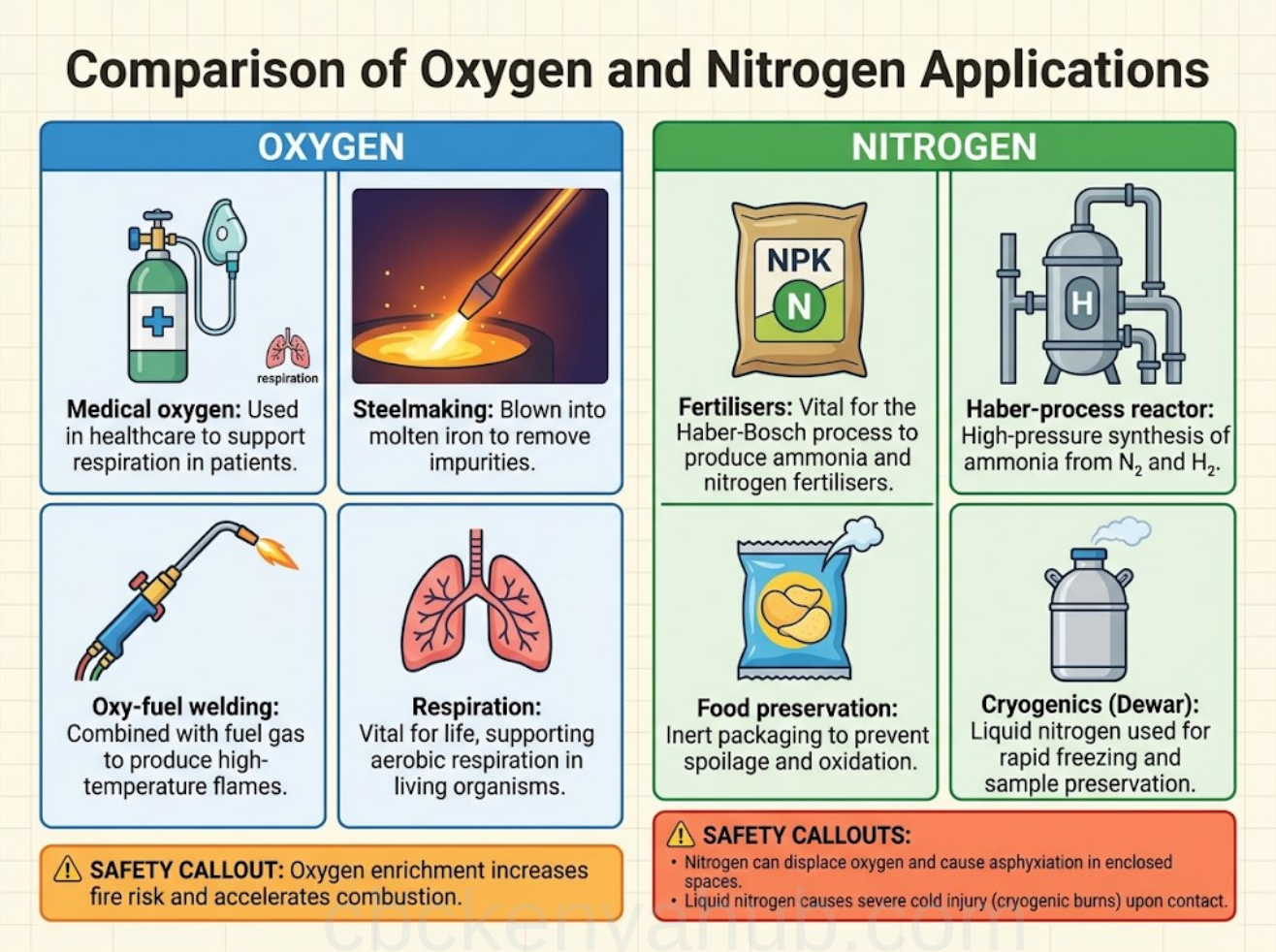
Uses of nitrogen

- Ammonia manufacture: nitrogen reacts with hydrogen in the Haber process.
- Fertiliser manufacture: nitrogen-containing compounds supply plants with nitrogen in usable forms.
- Food packaging: nitrogen can displace oxygen and reduce oxidation of some packaged foods.
- Inert atmospheres: nitrogen is used where limiting contact with oxygen or moisture is important, provided the process is designed safely.
- Liquid nitrogen: used for rapid cooling and cryogenic applications by trained personnel.

Liquid nitrogen is extremely cold and can cause severe cold burns. It can also displace oxygen in enclosed spaces, creating an asphyxiation hazard. It must never be stored in a sealed container not designed for cryogenic pressure, and it should only be handled using approved cryogenic equipment.



1.2.11 Compounds of Oxygen and Nitrogen

Oxygen and nitrogen form many compounds. Their properties depend on the other element present, the bonding, oxidation state and structure. Examples relevant to everyday life and industry include water (H_2O), carbon dioxide (CO_2), metal oxides, ammonia (NH_3), ammonium salts and nitrogen oxides.

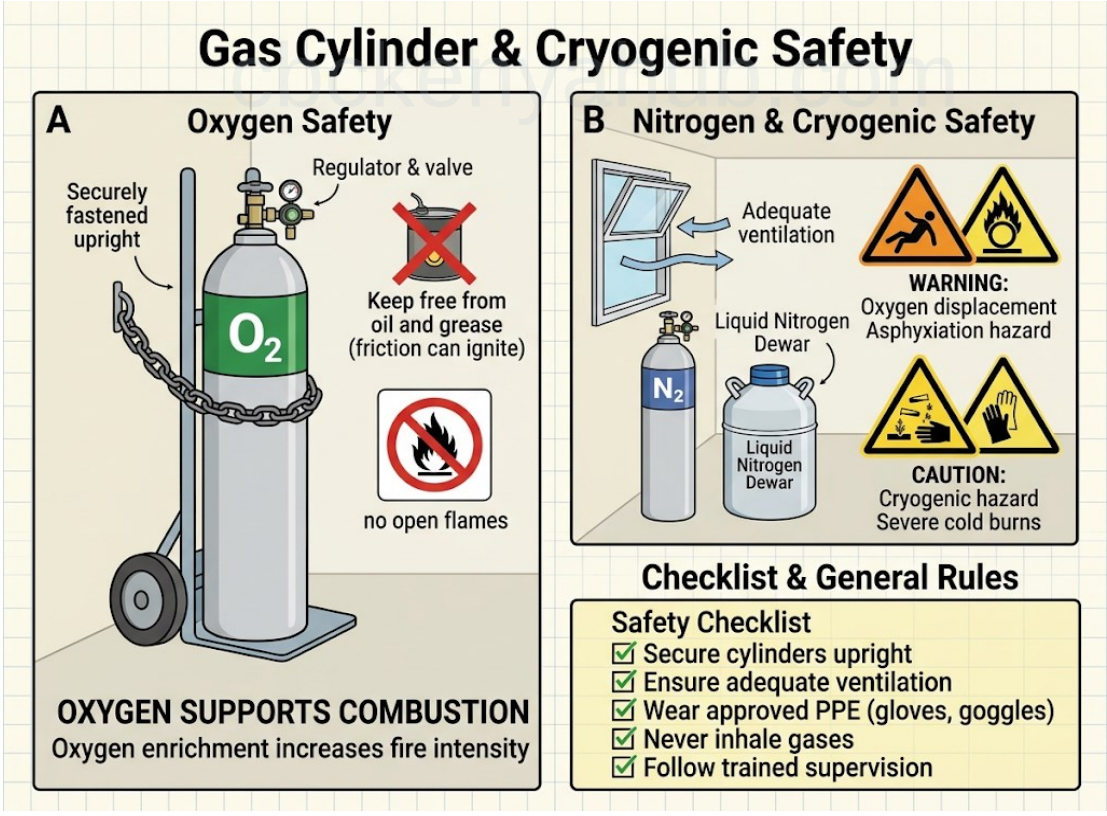
Compound	Key properties or role	Important note
Water, H_2O	Essential solvent and biological substance	Its properties differ greatly from hydrogen and oxygen gases
Carbon dioxide, CO_2	Produced by complete combustion and respiration; used in some fire extinguishers	Does not support ordinary combustion and can displace air in confined spaces
Ammonia, NH_3	Used in fertiliser manufacture; alkaline in water	Irritating and hazardous at sufficient concentration; never inhale directly
Ammonium salts	Many are soluble ionic compounds used in fertilisers	The ammonium ion is NH_4^+
Nitrogen oxides, e.g. NO_2	Produced in some high-temperature combustion and industrial processes	Contribute to air pollution and acid deposition

1.2.12 Environmental and Safety Considerations

Oxygen and nitrogen are natural components of air, but human activities can alter the amounts of other gases and produce harmful compounds. Nitrogen oxides formed during high-temperature combustion contribute to photochemical smog and acid deposition. Fertiliser runoff can add excess nitrate to water bodies, encouraging algal growth and reducing dissolved oxygen as the algae decompose. These effects demonstrate that the environmental impact of a substance depends on how it is produced, used and released.

Nitrogen gas is not toxic in the same way as a reactive poison, but it can be dangerous because it displaces oxygen. An atmosphere with insufficient oxygen can cause unconsciousness and death without obvious warning. Oxygen itself is not flammable, but oxygen enrichment makes many materials ignite more easily and burn more intensely.

- Keep oxygen cylinders secured and away from oil, grease, flames and ignition sources.
- Never use oxygen to blow dust from clothing or to freshen an enclosed room.
- Do not enter an enclosed space where nitrogen may have accumulated without an approved confined-space safety procedure.
- Use fume cupboards or appropriate ventilation for hazardous gases and follow school safety instructions.
- Dispose of chemicals according to teacher instructions; do not pour unknown reaction mixtures into drains.



1.2.13 Comparing Oxygen and Nitrogen in Experiments

A good practical comparison separates observations from inferences. An observation is what is directly seen, such as a glowing splint relighting. An inference is the explanation, such as the gas supports combustion and is consistent with oxygen. Students should avoid writing “oxygen burns” because the test shows that oxygen supports combustion, not that oxygen is a fuel.

Test / property	Oxygen	Nitrogen
Appearance	Colourless gas	Colourless gas
Glowing splint	Relights or glows more brightly	Does not relight the splint
Support of combustion	Supports combustion	Does not support ordinary combustion
Common atmospheric proportion	About 21% by volume	About 78% by volume
Important caution	Enrichment increases fire risk	Can displace oxygen and cause asphyxiation

No single negative test is enough to identify every unknown gas. If a gas does not relight a glowing splint, it might be nitrogen, carbon dioxide or another gas. A positive test and the preparation method, together with the other evidence, give a stronger identification.

1.2.14 Practical Skills: Planning a Gas Experiment

A practical investigation should have a clear aim, a labelled apparatus diagram, a list of chemicals and equipment, a method written in logical steps, expected observations, a conclusion and safety precautions. A good method specifies how gas is collected and how the experimenter knows when to stop.

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Example practical plan: Testing a sample of oxygen

Aim: To test whether a prepared gas sample has the characteristic property of supporting combustion.

Apparatus: A gas jar containing the sample, wooden splint, ignition source and appropriate eye protection.

Method: Under teacher supervision, prepare a glowing splint, remove the gas jar cover as instructed and introduce the splint carefully. Observe the response without inhaling gas or placing the face over the jar.

Expected observation: The glowing splint relights or glows more brightly.

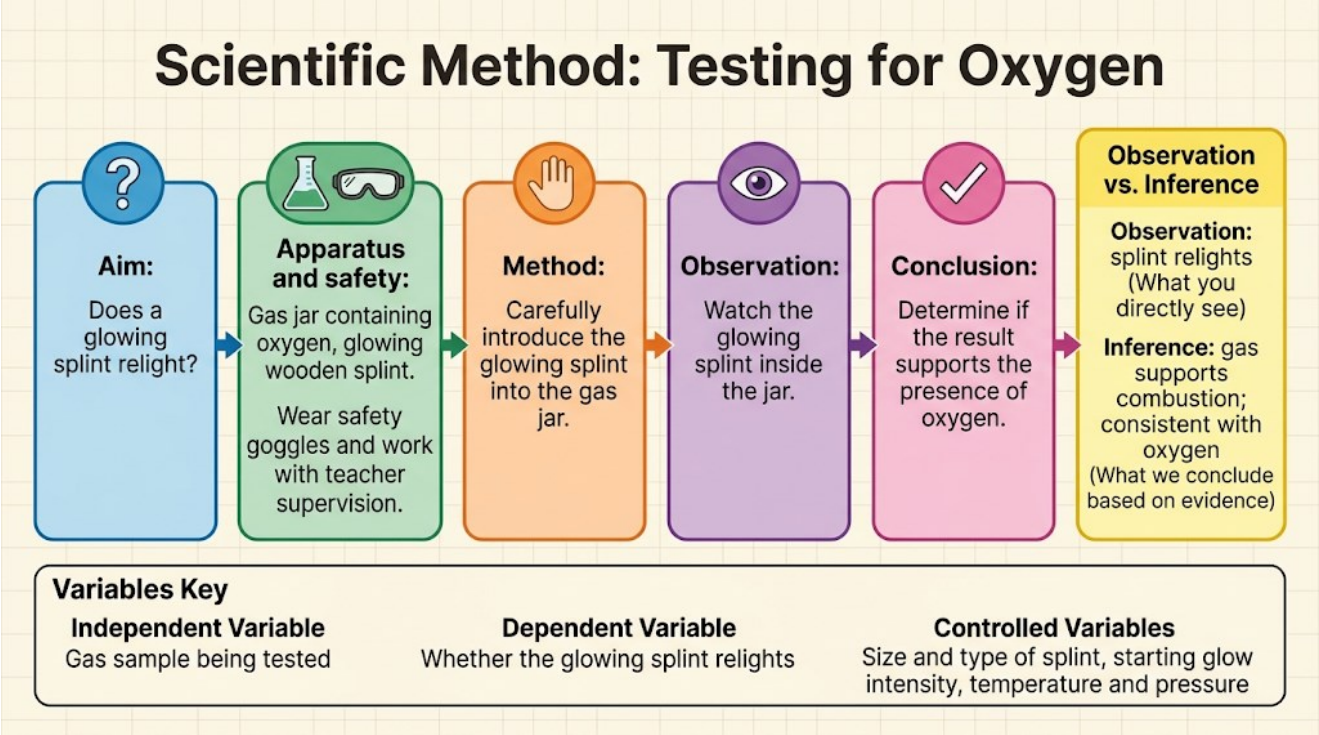
Conclusion: The observation is consistent with oxygen because oxygen supports combustion. The test should be interpreted alongside the preparation method and any other evidence.

Safety: Use only small quantities; keep combustible materials away from oxygen-enriched areas; handle hot splints carefully; and follow the teacher’s instructions.

Variables and fair testing

- Independent variable: the factor deliberately changed, such as the gas being tested.
- Dependent variable: the response observed, such as whether the glowing splint relights.

- Controlled variables: conditions kept consistent, such as the type and initial glow of the splint, the size of the sample and the procedure used.
- Repeat observations where appropriate and record them honestly.



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1.2.15 Summary: Non-metals

- Non-metals generally lack metallic conductivity, malleability and ductility, but exceptions and variations exist.
- Oxygen gas is O_2 and makes up about 21% of dry air by volume; nitrogen gas is N_2 and makes up about 78%.
- Oxygen is colourless, odourless, slightly soluble in water and supports combustion.
- Oxygen reacts with many elements to form oxides; the reaction may be rapid or slow.
- Laboratory oxygen can be prepared by catalytic decomposition of hydrogen peroxide and collected over water.
- Nitrogen is relatively unreactive at ordinary conditions because of the strong triple bond in N_2, but it reacts under suitable conditions.
- Nitrogen can be prepared in the laboratory by an approved method and collected over water because it is only slightly soluble.
- Industrial oxygen and nitrogen are separated from purified liquefied air by fractional distillation because their boiling points differ.

- Nitrogen is more volatile than oxygen because its boiling point is lower, so it is collected toward the top of the fractionating column.
- Oxygen is used in medicine, steelmaking and metal fabrication; nitrogen is used in ammonia manufacture, fertilisers, food packaging and inert atmospheres.
- Oxygen enrichment increases fire risk. Nitrogen can displace oxygen and cause asphyxiation; liquid nitrogen presents a severe cold hazard.
- Observations and inferences must be distinguished. A glowing splint relighting shows that a gas supports combustion, not that it is itself a fuel.

Self-check Questions

13. State three general properties of non-metals and give one exception to a generalisation.
14. Explain why oxygen is said to support combustion rather than burn itself.
15. Write the equation for the catalytic decomposition of hydrogen peroxide to form oxygen.
16. Explain why oxygen can be collected over water.
17. State two precautions needed when preparing and collecting a gas in the laboratory.
18. Explain why nitrogen is relatively unreactive at ordinary temperatures.
19. Describe the main stages in industrial separation of oxygen and nitrogen from air.
20. Explain why nitrogen is collected toward the top of a fractional distillation column.
21. Give three uses of oxygen and three uses of nitrogen.
22. Explain two safety hazards associated with oxygen and nitrogen.

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INTEGRATED REVISION AND EXAMINATION PRACTICE

A. Short-answer questions

23. Define the mole and state the value of the Avogadro constant.
24. Distinguish between relative formula mass and molar mass.
25. Calculate the molar mass of Al_2O_3 using Al = 27 and O = 16.
26. Calculate the amount of substance in 7.30 g of HCl using H = 1 and Cl = 35.5.
27. How many oxygen atoms are present in 0.40 mol of $Al_2(SO_4)_3$?
28. Calculate the concentration of a solution containing 0.0150 mol in 75.0 cm^3.
29. State two reasons why a standard solution may have a concentration different from the intended value.
30. Explain why a conical flask is rinsed with distilled water rather than with the analyte solution between titrations.
31. Write a balanced equation for the complete neutralisation of H_2SO_4 by KOH.
32. State the role of acid in the reaction between manganate(VII) and iron(II) ions.
33. Explain the difference between a limiting reactant and an excess reactant.
34. State the conditions under which gas-volume ratios can be compared directly using equation coefficients.
35. Give the approximate proportions of nitrogen and oxygen in dry air.
36. State the test for oxygen and the expected observation.
37. Explain why water and carbon dioxide are removed before air is liquefied industrially.

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B. Structured calculation practice

Question 1. A student dissolves 5.30 g of anhydrous sodium carbonate, Na_2CO_3, and makes the solution up to 500 cm^3. Calculate (a) its molar mass, (b) the amount dissolved, and (c) the concentration.

Question 2. 25.0 cm^3 of sodium hydroxide is neutralised by 18.60 cm^3 of 0.150 mol dm^{-3} hydrochloric acid. Calculate the concentration of the sodium hydroxide.

Question 3. 2.40 g of magnesium burns completely in oxygen. Calculate the amount of magnesium, the amount of magnesium oxide produced and the mass of magnesium oxide. Use Mg = 24 and O = 16.

Question 4. A gas sample occupies 480 cm^3 at conditions where the molar gas volume is 24.0 $dm^3\ mol^{-1}$. Calculate the amount of gas.

Question 5. Nitrogen reacts with hydrogen to form ammonia according to $N_2 + 3H_2 \rightarrow 2NH_3$. If 90 cm^3 of hydrogen reacts completely with excess nitrogen, calculate the volume of ammonia formed at the same temperature and pressure.

C. Answers and marking guidance

1. The mole is the amount of substance containing exactly $6.02214076 \times 10^{23}$ specified entities; commonly rounded to 6.02×10^{23} mol^{-1}.

2. Relative formula mass is a dimensionless sum of relative atomic masses; molar mass is the mass of one mole and is expressed in g mol^{-1}.

3. $M(Al_2O_3) = 2(27) + 3(16) = 102$ g mol^{-1}.

4. $M(HCl) = 36.5$ g mol^{-1}; $n = 7.30/36.5 = 0.200$ mol.

5. One formula unit $Al_2(SO_4)_3$ contains 12 oxygen atoms. Amount of oxygen atoms = $0.40 \times 12 = 4.8$ mol oxygen atoms.

6. Volume = 0.0750 dm^3; $c = 0.0150/0.0750 = 0.200$ mol dm^{-3}.

7. Possible reasons: incomplete transfer; inaccurate weighing; making up to the wrong mark; contamination; or not mixing thoroughly.

8. Distilled water does not add analyte and therefore does not change the amount transferred by the pipette; residual analyte would add an unknown amount.

9. $H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$.

10. Acid supplies H^+ ions and establishes the conditions under which manganate(VII) is reduced to Mn^{2+} in the standard reaction.

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11. The limiting reactant is used up first and determines the theoretical amount of product; an excess reactant remains after the limiting reactant is consumed.

12. The gases must be at the same temperature and pressure, and the balanced equation must represent the reaction being considered.

13. Nitrogen ≈78%; oxygen ≈21% by volume in dry air.

14. Introduce a glowing splint; it relights or glows more brightly.

15. Water and carbon dioxide can freeze or solidify at cryogenic temperatures and block equipment.

Structured practice: Q1: M = 106 g mol^{-1}; $n = 5.30/106 = 0.0500$ mol; $c = 0.0500/0.500 = 0.100$ mol dm^{-3}. Q2: moles HCl = $0.150 \times 0.01860 = 0.00279$ mol; 1:1 ratio gives the same NaOH amount; c(NaOH) = $0.00279/0.0250 = 0.1116$ mol $dm^{-3} \approx 0.112$ mol dm^{-3}. Q3: $n(Mg) = 2.40/24 = 0.100$ mol; ratio Mg:MgO = 1:1; $n(MgO) = 0.100$ mol; $M(MgO) = 40$ g mol^{-1}; mass = 4.00 g. Q4: 480 $cm^3 = 0.480$ dm^3; $n = 0.480/24.0 = 0.0200$ mol. Q5: $H_2:NH_3 = 3:2$; NH_3 volume = $90 \times 2/3 = 60$ cm^3.

Glossary of Key Terms

Term	Meaning
Aliquot	An accurately measured portion of a solution.
Amount of substance	A measure of the number of specified entities, expressed in moles.
Analyte	The solution being analysed in a titration.
Avogadro constant	The number of specified entities per mole, approximately 6.02×10^{23} mol^{-1}.
Balanced equation	A chemical equation with equal numbers of each type of atom on both sides.
Catalyst	A substance that changes reaction rate without being consumed overall.
Concentration	Amount of solute per volume of solution; often expressed in mol dm^{-3}.
Concordant titres	Titre measurements that agree closely with each other.
Empirical formula	The simplest whole-number ratio of atoms in a compound.
Endpoint	The observed point in a titration at which the indicator changes as specified.
Limiting reactant	The reactant that is consumed first and limits product formation.
Molar gas volume	The volume occupied by one mole of gas at specified temperature and pressure.
Molar mass	Mass of one mole of a substance, expressed in g mol^{-1}.
Mole ratio	The ratio of amounts indicated by coefficients in a balanced equation.
Oxidation	Loss of electrons; in many introductory reactions, also described as addition of oxygen.
Redox reaction	A reaction involving oxidation and reduction together.
Relative atomic mass	The weighted average mass of atoms of an element relative to one-twelfth of a carbon-12 atom.
Relative formula mass	The sum of relative atomic masses represented in a formula.
Standard solution	A solution whose concentration is accurately known.
Titrant	The solution delivered from a burette during titration.
Titre	The volume delivered from a burette, calculated from final minus initial reading.

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Final Revision Checklist

- I can explain what the mole measures and use the Avogadro constant.
- I can calculate relative formula masses and molar masses from chemical formulas.
- I can convert between mass, amount, number of particles, concentration and gas volume.
- I can prepare a standard solution and explain why quantitative transfer matters.
- I can interpret titration apparatus, calculate titres and solve acid–base titration problems.
- I can use balanced redox equations to calculate amounts and concentrations.
- I can apply mole ratios, identify limiting reactants and calculate theoretical yield.
- I can describe oxygen and nitrogen, their tests, laboratory preparation and industrial separation.
- I can explain common uses and safety precautions for oxygen, nitrogen and their compounds.

- I can distinguish direct observations from scientific inferences.

End of Strand 1.0: Inorganic Chemistry — Grade 11 Chemistry,

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About These Notes

These notes follow Strand 2.0: Physical Chemistry in the supplied Grade 11 Chemistry Curriculum Design (KICD, October 2025). The strand is organised into three sub-strands: 2.1 Salts, 2.2 Gas Laws and 2.3 Reaction Rates. The explanations are written in connected textbook-style prose for revision, supported by scan-friendly bullet points, worked examples, practical guidance, summaries and self-check questions.

How to Use the Notes

- Read the explanation first, then study any equations and note the units used.
- Attempt each worked example before reading the solution steps.
- For practical work, follow teacher instructions, use suitable personal protective equipment and never mix chemicals without approval.
- Use the end-of-section summaries and revision questions to check understanding.
- Keep gas-law temperatures in kelvin and show substitutions and units in every calculation.

Strand Map

Sub-strand	Main learning focus
2.1 Salts	Nature and classification of salts; solubility; preparation; crystallisation; salt reactions; ionic equations; uses and safety.
2.2 Gas Laws	Gas pressure and particle model; Boyle's, Charles's, pressure and combined gas laws; Avogadro's law; ideal gas equation; applications.
2.3 Reaction Rates	Meaning and measurement of rate; collision theory; concentration, pressure, temperature, surface area and catalysts; graphs and practical investigations.

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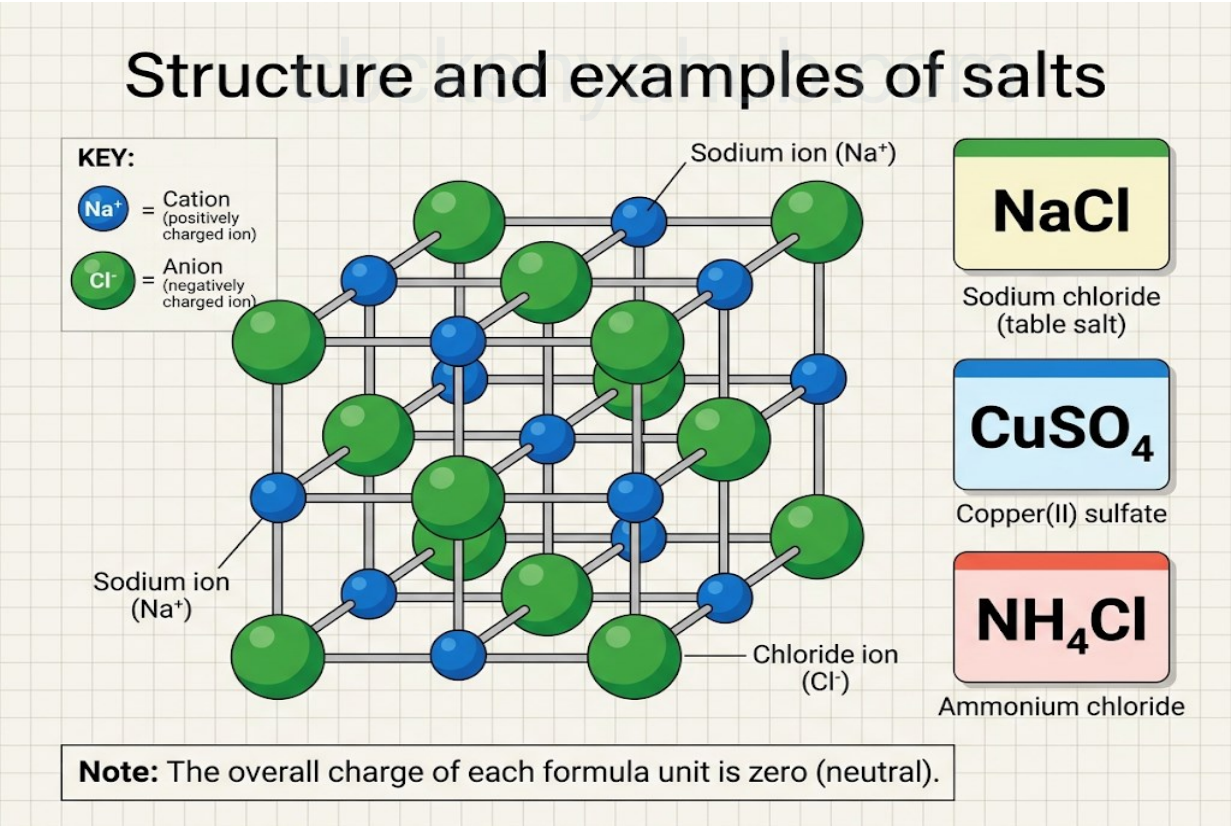
SUB-STRAND 2.1: SALTS

Curriculum focus: describe and investigate salts, their preparation, solubility and reactions; interpret ionic equations; and connect the chemistry of salts to familiar laboratory, environmental and industrial contexts.

1. What Is a Salt?

A salt is an ionic compound formed when the replaceable hydrogen ion or ions of an acid are wholly or partly replaced by metal ions or ammonium ions. Salts contain positive ions (cations) and negative ions (anions) arranged in an electrically neutral lattice when solid. In water, soluble salts separate into mobile ions, which explains why their aqueous solutions can conduct electricity.

- Examples include sodium chloride (NaCl), copper(II) sulfate ($CuSO_4$), ammonium chloride (NH_4Cl), calcium carbonate ($CaCO_3$) and potassium nitrate (KNO_3).
- A salt does not have to be neutral in pH. Some salts produce acidic or alkaline solutions because their ions react with water; the behaviour depends on the ions present.
- Salts are different from common table salt. Sodium chloride is only one member of the much larger chemical family called salts.



2. Naming and Writing Formulae of Salts

To name a salt, identify its positive ion first and its negative ion second. The positive ion is often a metal ion or ammonium, NH_4^+. The negative ion may be a single-element ion, such as chloride, or a polyatomic ion, such as sulfate or nitrate. Where a metal forms ions with different charges, Roman numerals identify the charge in the name.

- Common anions: chloride Cl^-, bromide Br^-, iodide I^-, oxide O^{2-}, sulfide S^{2-}, nitrate NO_3^-, sulfate SO_4^{2-}, carbonate CO_3^{2-} and phosphate PO_4^{3-}.
- Formulae must be electrically neutral. Choose the smallest whole-number ratio of ions that makes total positive charge equal total negative charge.
- Use brackets around a polyatomic ion when more than one of that ion is needed, for example $Ca(NO_3)_2$.

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- Use brackets around a polyatomic ion when more than one of that ion is needed, for example $Ca(NO_3)_2$.

Worked Example 1: Formula of aluminium sulfate

Problem. Write the formula of aluminium sulfate.

Step 1: Aluminium forms Al^{3+}; sulfate is SO_4^{2-}.

Step 2: Two Al^{3+} ions give +6 and three sulfate ions give −6.

Answer. The formula is $Al_2(SO_4)_3$.

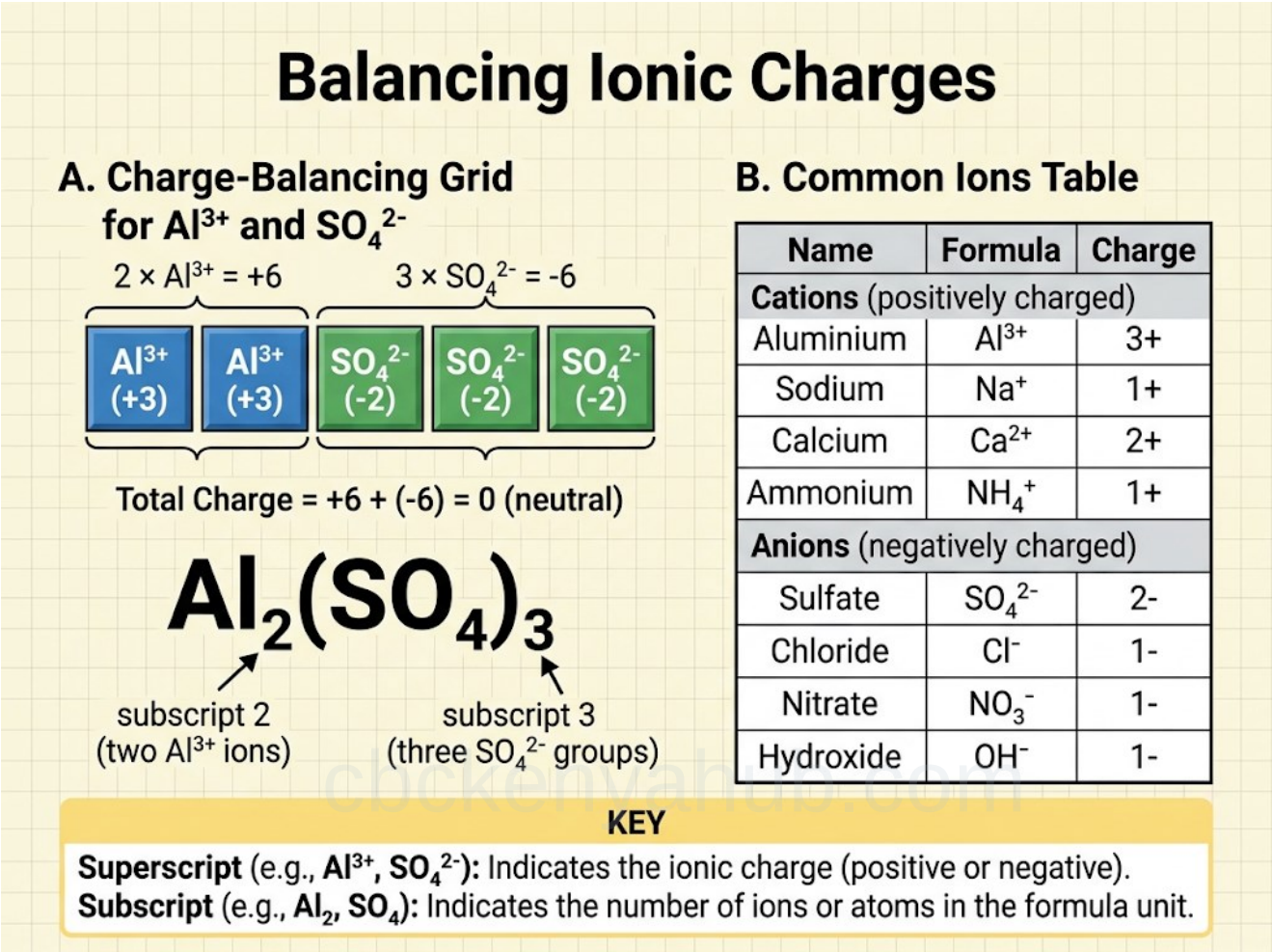
Worked Example 2: Name $FeCl_3$

Problem. Name $FeCl_3$.

Step 1: Each chloride ion has charge −1, so three chloride ions contribute −3.

Step 2: The iron ion must therefore be Fe^{3+}; use the Roman numeral III.

Answer. The name is iron(III) chloride.



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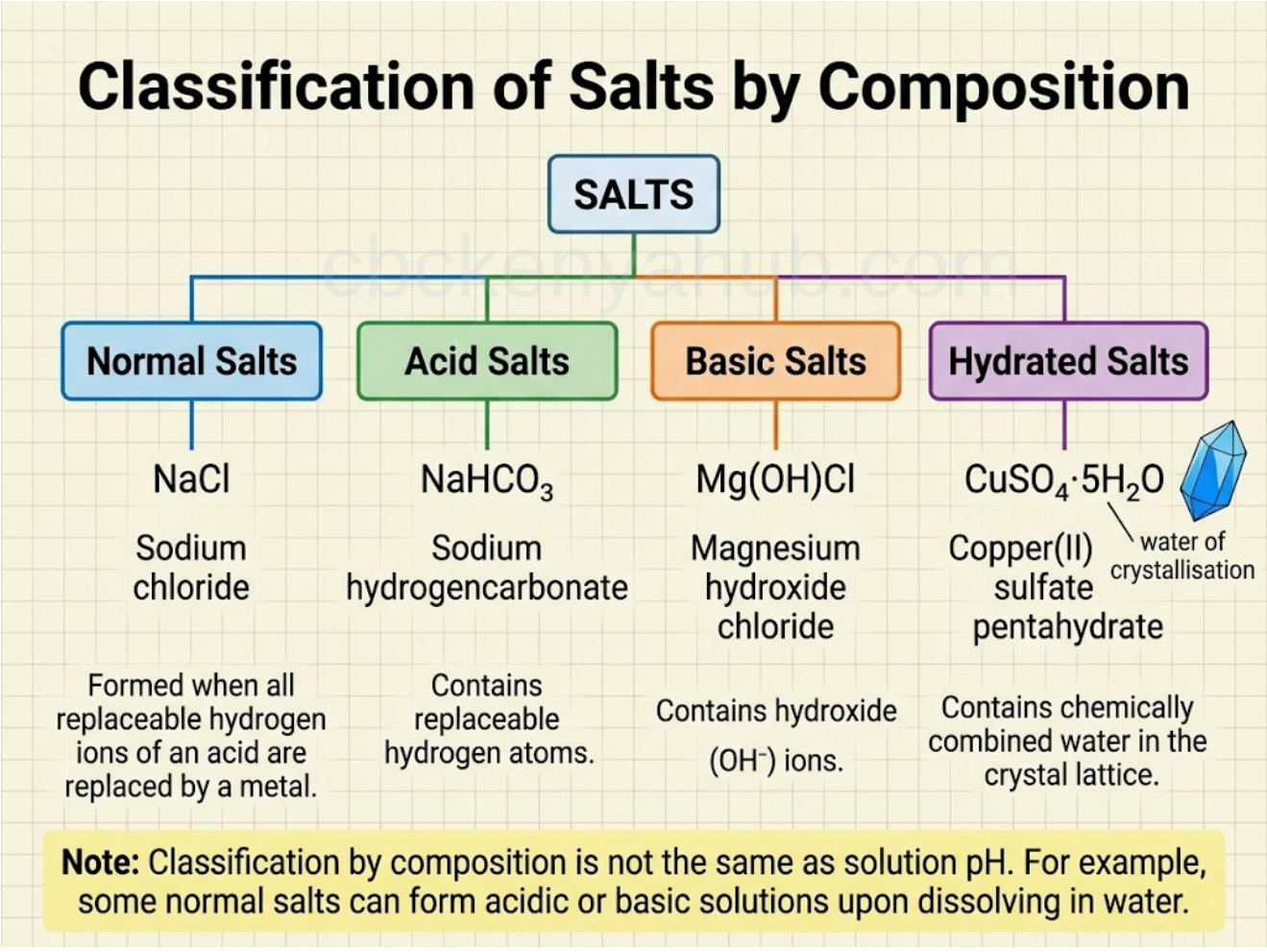
3. Classification of Salts

Salts may be classified by their composition and by how they behave in water. Normal salts form when all replaceable acidic hydrogen ions are replaced. Acid salts retain hydrogen in the anion because only partial replacement has occurred in a polybasic acid. Basic salts contain hydroxide groups alongside other anions. The term “acid salt” describes composition and does not automatically mean the solution has pH below 7.

- Normal salts include Na_2SO_4, KNO_3 and $CaCl_2$.
- Acid salts include $NaHCO_3$ (sodium hydrogencarbonate) and $NaHSO_4$ (sodium hydrogensulfate).
- Double salts contain two different cations or anions in a single crystalline compound, for example potash alum, $KAl(SO_4)_2 \cdot 12H_2O$.
- Hydrated salts contain water of crystallisation within the crystal structure, e.g. $CuSO_4 \cdot 5H_2O$. An anhydrous salt lacks this water.



4. Solubility of Salts

Solubility describes the maximum amount of a solute that dissolves in a given amount of solvent at a stated temperature. A soluble salt produces a solution when mixed with enough water; a sparingly soluble or insoluble salt may leave solid particles. Solubility rules help predict whether a precipitate forms when two aqueous solutions are mixed.

- All common sodium, potassium and ammonium salts are soluble in water.
- All common nitrates are soluble.
- Most chlorides are soluble, but silver chloride and lead(II) chloride are important exceptions.
- Most sulfates are soluble; barium sulfate and lead(II) sulfate are insoluble, while calcium sulfate is sparingly soluble.
- Most carbonates are insoluble except those of sodium, potassium and ammonium. These are school-level rules; use the supplied solubility table for unfamiliar salts.

Worked Example 3: Predicting a precipitate

Problem. Aqueous silver nitrate is mixed with aqueous sodium chloride. Predict the products and observation.

Step 1: Exchange the ions: possible products are AgCl and $NaNO_3$.

Step 2: Nitrates and sodium salts are soluble; AgCl is insoluble.

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Answer. A white precipitate of AgCl forms: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$.

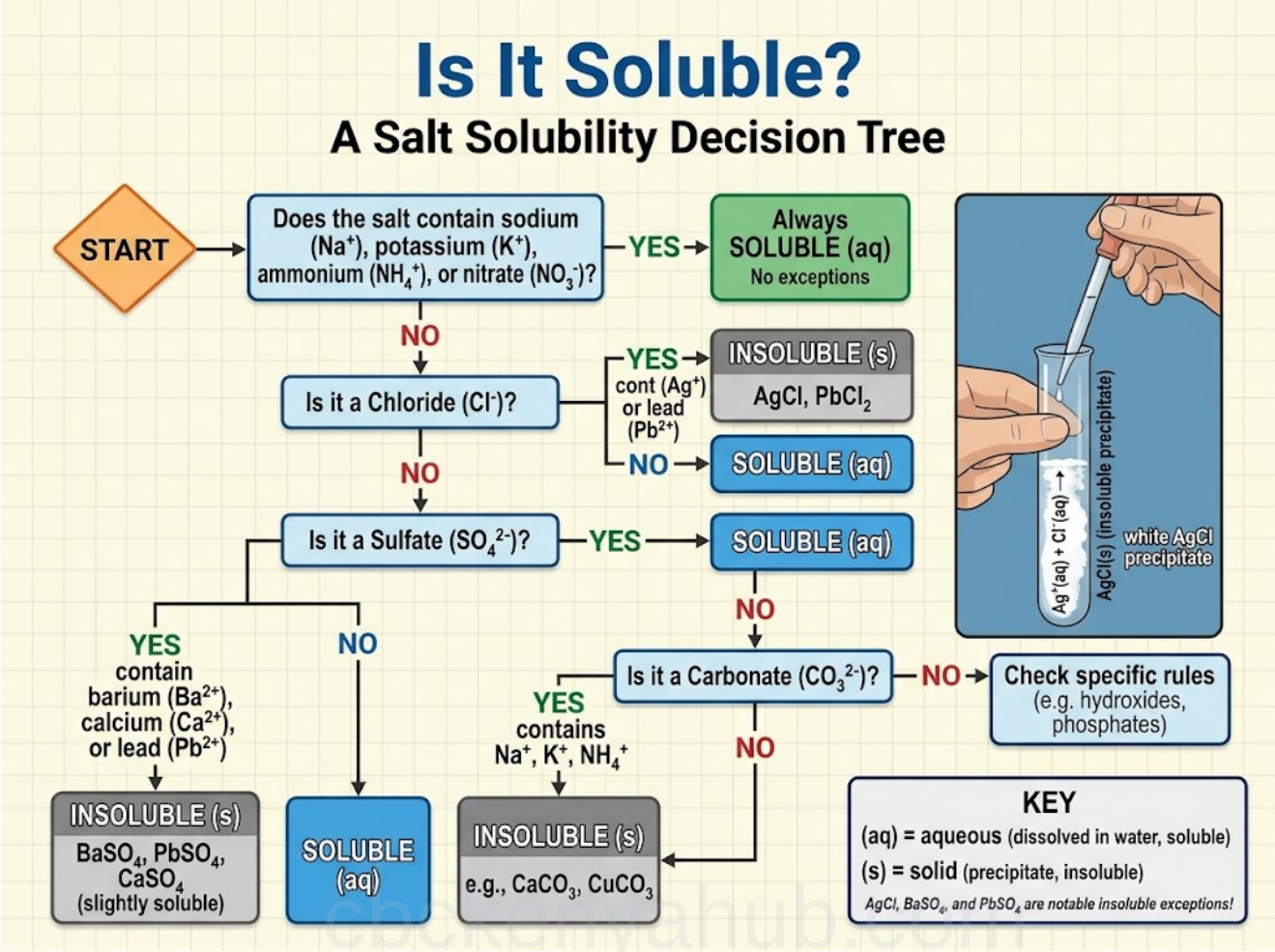
Worked Example 4: Potassium carbonate

Problem. Predict whether K_2CO_3 dissolves in water.

Step 1: Potassium salts are soluble, including potassium carbonate.

Step 2: The rule for potassium salts applies even though most carbonates are insoluble.

Answer. K_2CO_3 is soluble in water.



5. Preparing Soluble Salts from an Insoluble Base

An insoluble metal oxide or hydroxide can be reacted with a suitable dilute acid to prepare a soluble salt. The insoluble solid is added in excess so that all the acid is used up. The remaining solid is removed by filtration, leaving a salt solution that can be concentrated and crystallised.

- General pattern: acid + metal oxide → salt + water; acid + metal hydroxide → salt + water.
- Warm the dilute acid gently if instructed, then add the insoluble base in small portions while stirring.
- Continue adding base until a little remains undissolved; this indicates that the acid has been neutralised and base is in excess.
- Filter to remove excess solid. Heat the filtrate gently to concentrate it, then allow it to cool so crystals form.

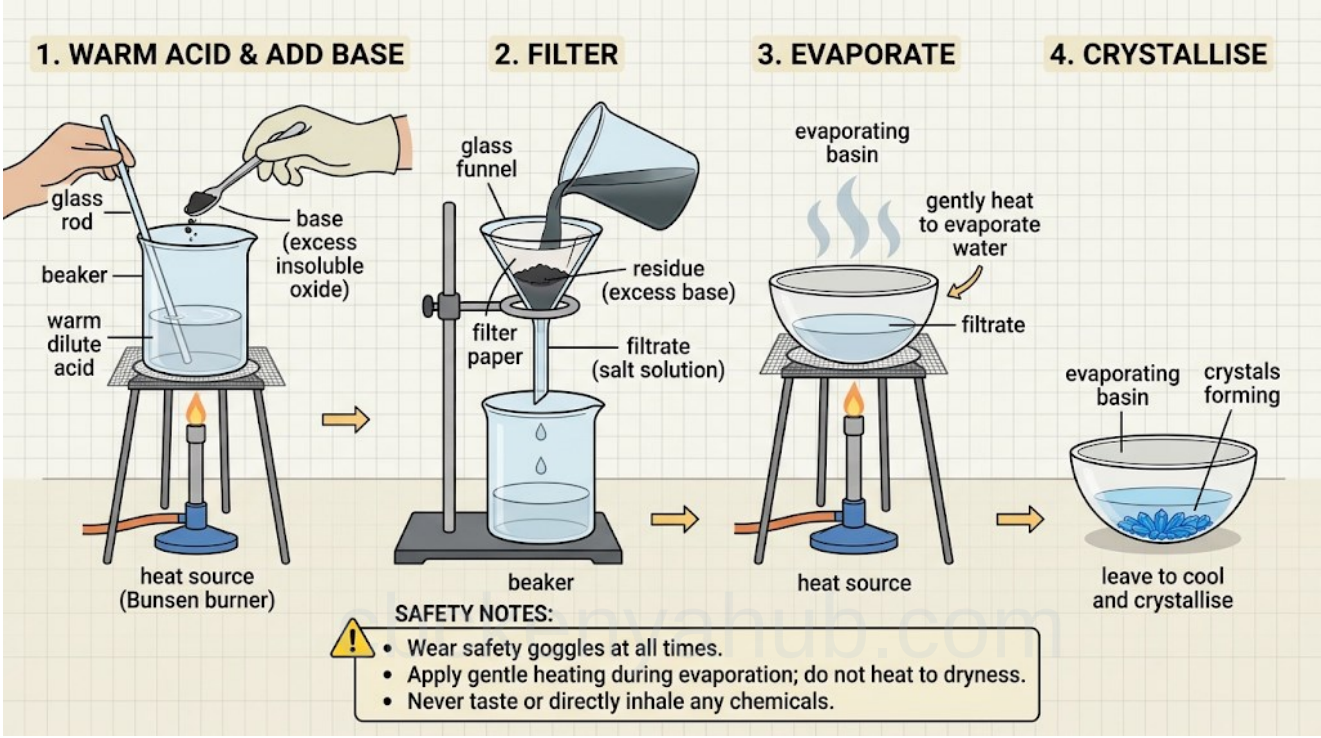
Worked Example 5: Copper(II) sulfate crystals

Problem. Describe how to prepare hydrated copper(II) sulfate crystals from dilute sulfuric acid and copper(II) oxide.

Step 1: Warm dilute sulfuric acid. Add black CuO in small portions while stirring until some remains.

Step 2: Filter off excess CuO. Gently evaporate part of the water from the blue filtrate, then cool it.

Answer. Blue $CuSO_4 \cdot 5H_2O$ crystals form. Equation: $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$.



6. Preparing Soluble Salts by Titration

When both acid and base are soluble, neither reactant can simply be removed by filtration. Titration determines the exact volumes that neutralise each other. First use an indicator to locate the end-point; then repeat using the same measured volumes but without indicator so that the product is not contaminated by indicator.

- Pipette a measured volume of alkali into a conical flask and add a few drops of a suitable indicator.
- Fill a burette with acid, record the initial reading, and add acid while swirling until the indicator changes colour at the end-point.
- Repeat until concordant titres are obtained. Use the mean titre.
- Repeat the neutralisation with the same volumes but without indicator, then crystallise the salt solution.

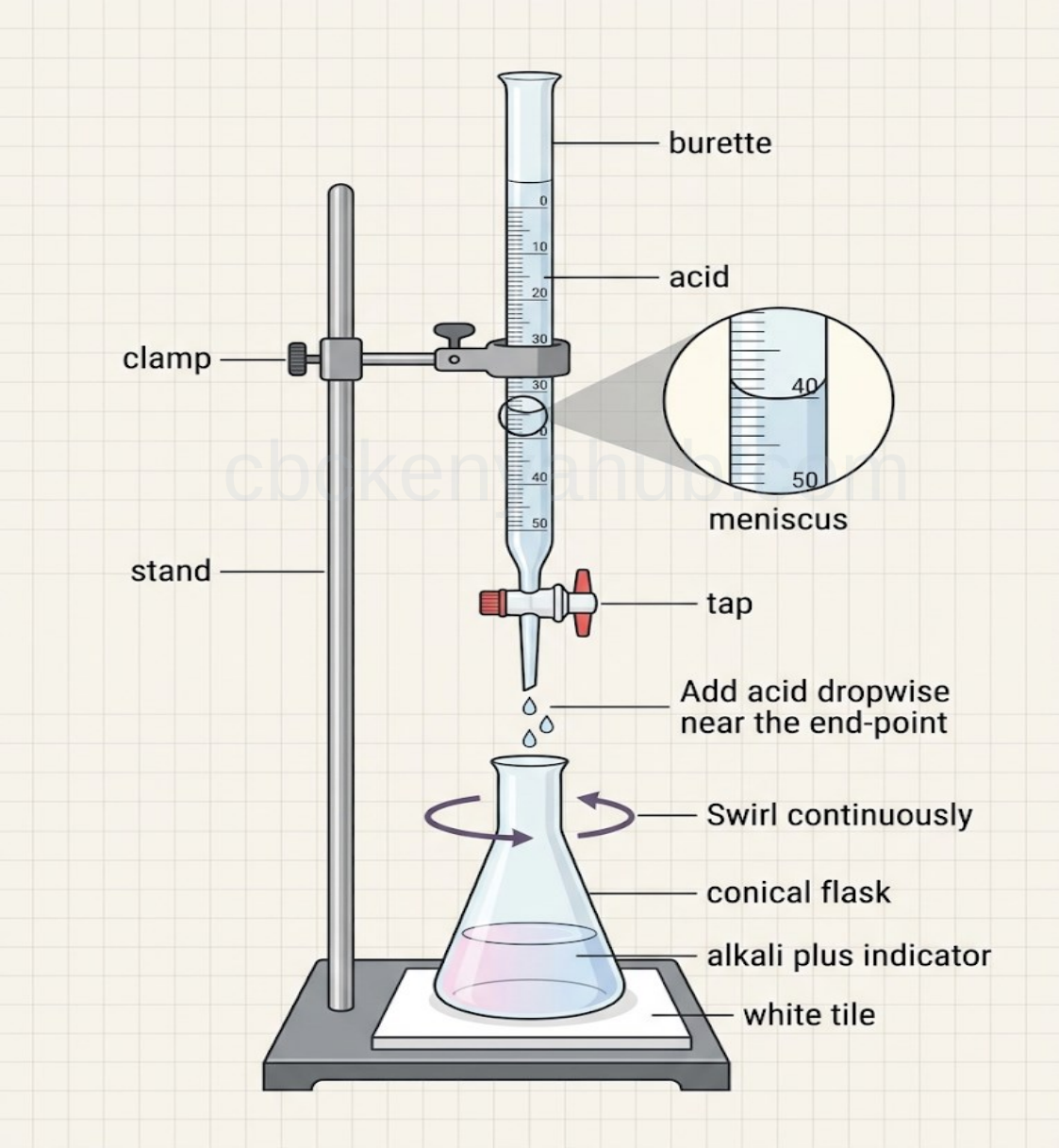
Worked Example 6: Titration volumes

Problem. A neutralisation requires 25.0 cm³ sodium hydroxide and a mean of 20.0 cm³ hydrochloric acid. What acid volume should be used for an indicator-free preparation?

Step 1: Use the same measured volume that gave neutralisation in the titration.

Step 2: The end-point volume of hydrochloric acid is 20.0 cm³.

Answer. Use 20.0 cm³ of hydrochloric acid with 25.0 cm³ sodium hydroxide, assuming the same stock solutions.



7. Preparing Insoluble Salts by Precipitation

An insoluble salt is often prepared by precipitation: two soluble salt solutions are mixed so that their ions combine to form an insoluble product. The solid is separated by filtration, washed with distilled water to remove soluble impurities and dried appropriately.

- Choose reactants that supply the required cation and anion in soluble form.
- Mix solutions while stirring. The appearance of a solid indicates precipitation.
- Filter the mixture, wash the precipitate with small portions of distilled water and dry it appropriately.
- Use ionic equations to show only the ions that take part in forming the precipitate.

Worked Example 7: Preparing barium sulfate

Problem. Write the equation for forming barium sulfate from barium chloride and sodium sulfate.

Step 1: The required ions are Ba^{2+} and SO_4^{2-}.

Step 2: Na^+ and Cl^- remain dissolved as sodium chloride.

Answer. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$; net ionic: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$.

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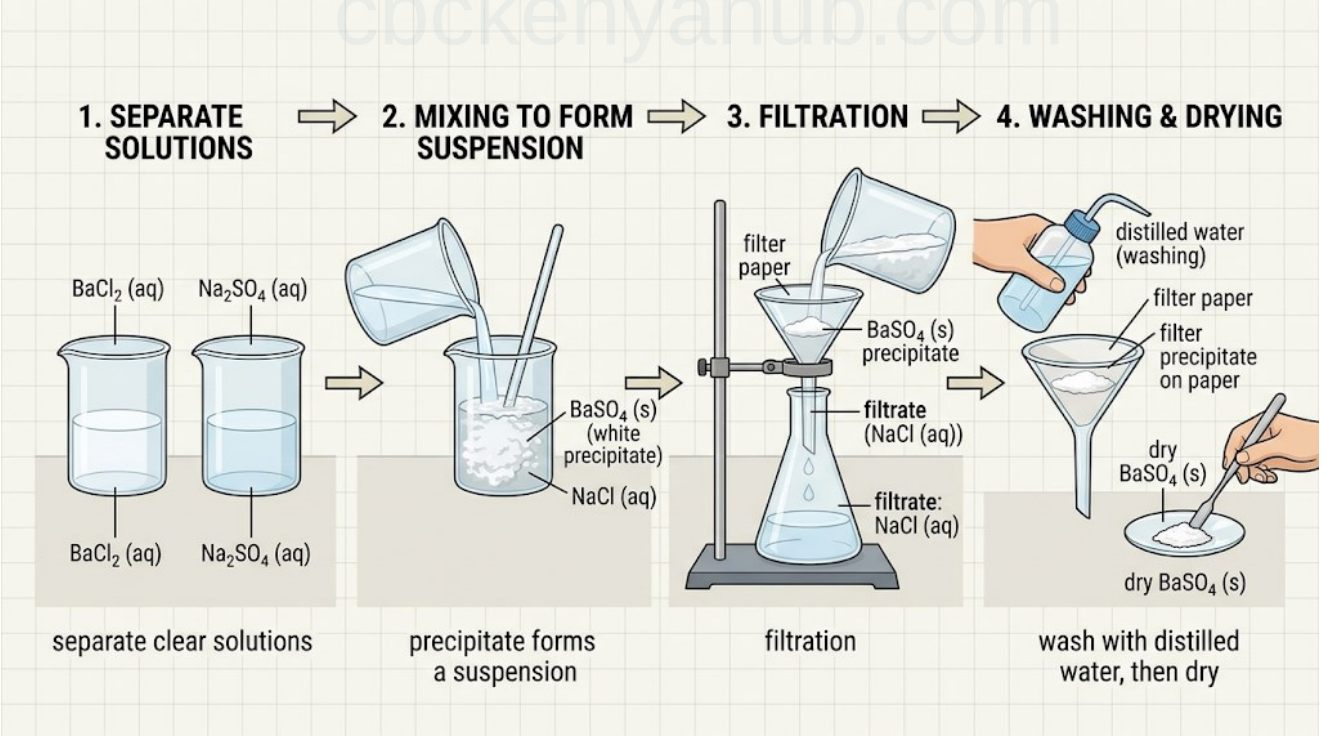

8. Crystallisation and Water of Crystallisation

Crystallisation produces a purified crystalline solid from a solution. The solution is concentrated until close to saturation, then allowed to cool. As it cools, the solubility of many salts decreases and crystals grow. The correct drying method depends on whether the salt is hydrated or anhydrous.

- Evaporate solvent gently to avoid spitting and decomposition.
- Allow the concentrated solution to cool slowly; slow cooling often produces larger crystals.
- Separate crystals from the remaining solution and dry carefully.
- Water of crystallisation is water incorporated in a definite ratio into a crystal structure, shown by a dot in $CuSO_4 \cdot 5H_2O$.

Worked Example 8: Hydrated copper sulfate

Problem. What does $CuSO_4 \cdot 5H_2O$ tell you?

Step 1: The dot separates the salt formula from water of crystallisation.

Step 2: The formula shows five water molecules per formula unit of copper(II) sulfate in the crystal.

Answer. Each formula unit contains $CuSO_4$ and five waters of crystallisation.

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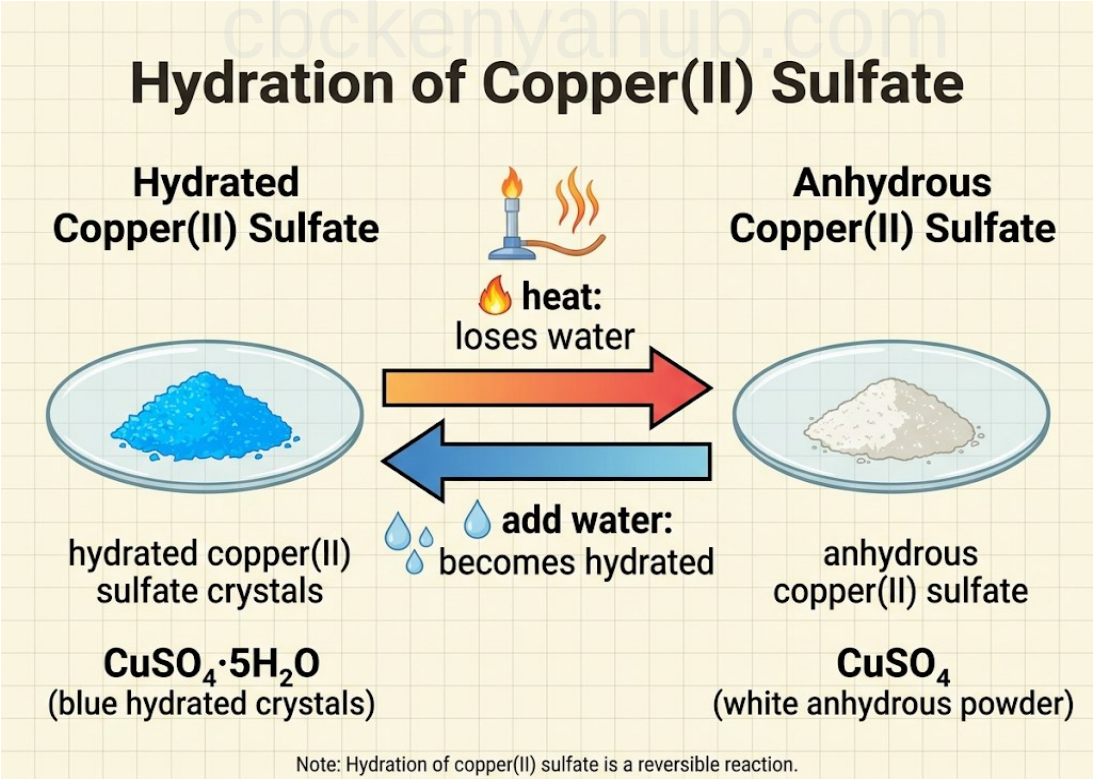


9. Reactions and Uses of Salts

Salts participate in precipitation, reactions with acids, thermal decomposition and displacement reactions. They are used in food preservation, fertilisers, water treatment, construction, medicine, industry and laboratory analysis.

- Carbonates react with dilute acids to produce a salt, water and carbon dioxide.
- Many metal carbonates decompose on strong heating to form a metal oxide and carbon dioxide.
- Some salts identify ions through characteristic precipitates; tests must be interpreted carefully.
- Fertilisers supply nutrients such as nitrogen, phosphorus and potassium; incorrect use can cause runoff and eutrophication.

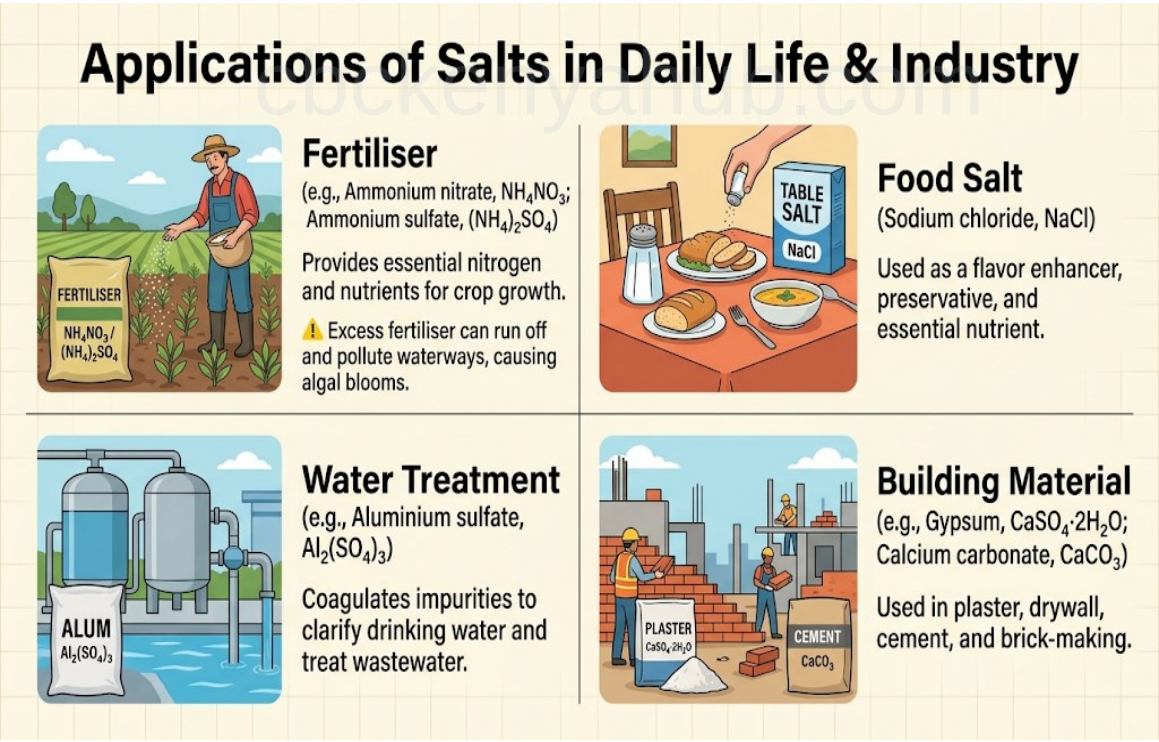
Worked Example 9: Acid and carbonate

Problem. Write the products when calcium carbonate reacts with hydrochloric acid.

Step 1: Acid + carbonate gives salt, water and carbon dioxide.

Step 2: The salt from calcium ions and chloride ions is calcium chloride.

Answer. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$.



10. Ionic and Net Ionic Equations

A complete ionic equation represents soluble ionic compounds in aqueous solution as separated ions. Ions unchanged on both sides are spectator ions. A net ionic equation removes these spectator ions and shows only the species undergoing chemical change.

- Split strong, soluble aqueous ionic compounds into ions; do not split solids, liquids, gases or weakly ionised substances.
- Balance atoms and total charge.
- Cancel ions that appear unchanged on both sides.
- Check that atoms and charge are conserved.

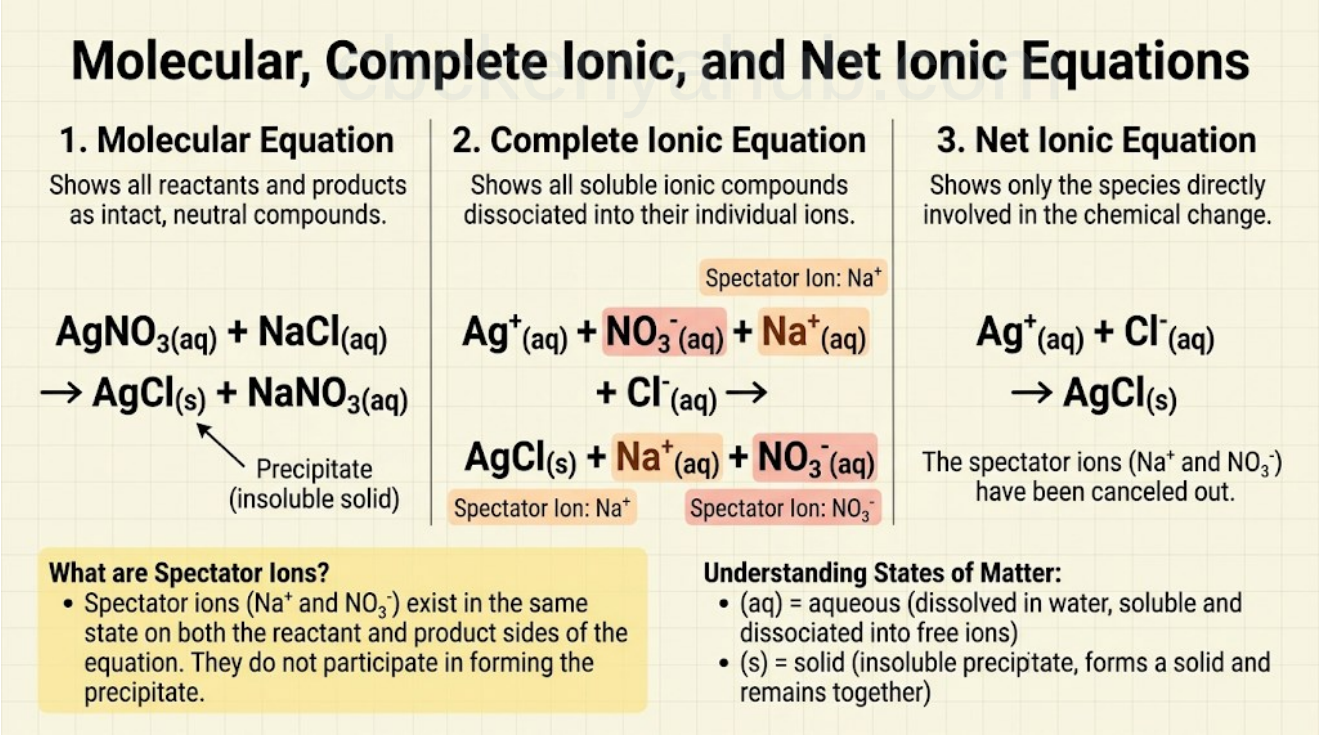
Worked Example 10: Silver chloride

Problem. Obtain the net ionic equation for $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$.

Step 1: Write ions: $Ag^+ + NO_3^- + Na^+ + Cl^- \rightarrow AgCl(s) + Na^+ + NO_3^-$.

Step 2: Na^+ and NO_3^- are spectator ions and cancel.

Answer. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.



Summary: Salts

- Salts are ionic compounds whose formulae have zero net charge.
- Solubility rules predict whether salts dissolve or precipitate.

- Soluble salts can be prepared using an excess insoluble base or by titration.
- Insoluble salts are prepared by precipitation, then filtered, washed and dried.
- Crystallisation purifies salts; hydrated salts contain water of crystallisation.
- Net ionic equations omit spectator ions.

Self-check Questions: Salts

- Write formulae for calcium nitrate, ammonium sulfate and aluminium chloride.
- Explain why excess copper(II) oxide is used when preparing copper(II) sulfate.
- Describe preparation of a pure insoluble salt by precipitation.
- Write molecular and net ionic equations for barium sulfate formation.
- State two differences between crystallisation and precipitation.
- Explain why a salt solution may be acidic or alkaline.

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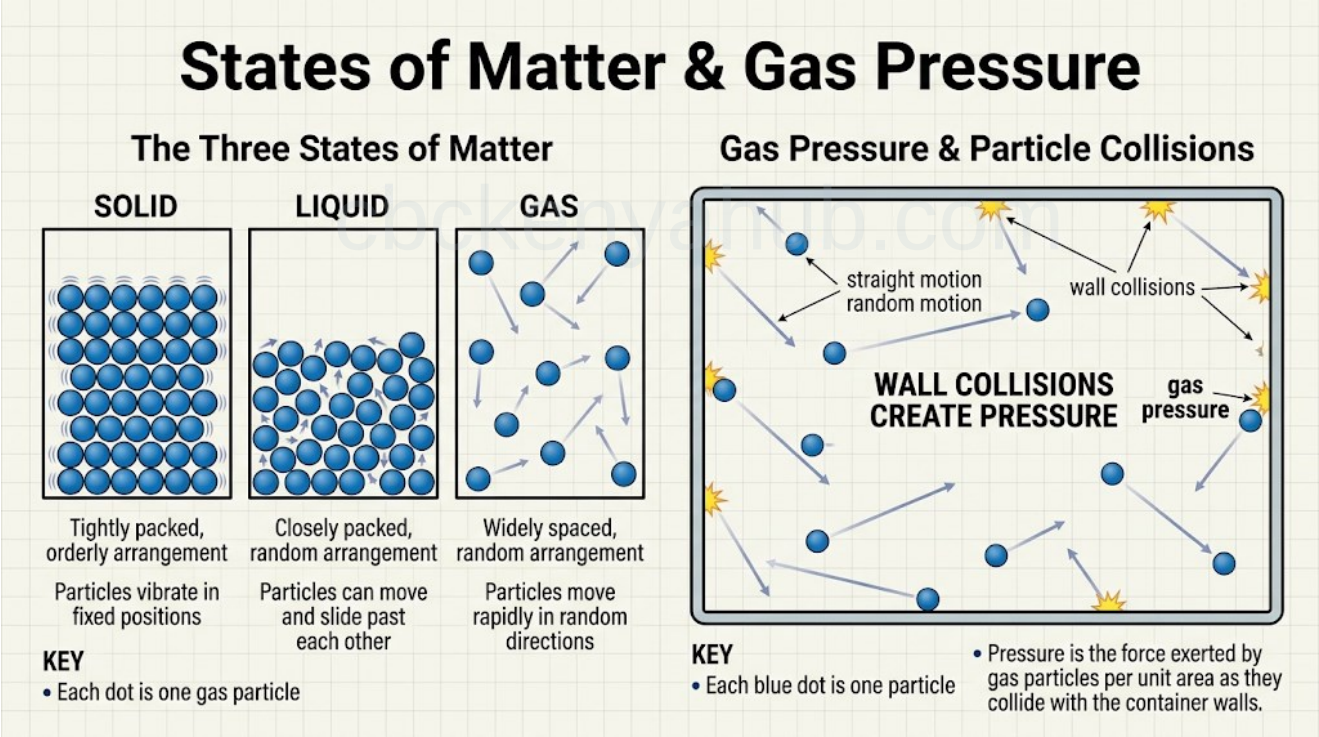
SUB-STRAND 2.2: GAS LAWS

Curriculum focus: explain the behaviour of gases under different conditions, apply gas laws to quantitative problems and appreciate their everyday and industrial applications.

1. Particle Model of Gases

A gas consists of particles moving rapidly and randomly in all directions. They are far apart relative to their own size, so gases are compressible. Collisions with container walls produce pressure. Temperature in kelvin is related to average kinetic energy. The ideal gas model assumes negligible particle volume and negligible intermolecular attraction; real gases approach it most closely at relatively low pressure and high temperature.

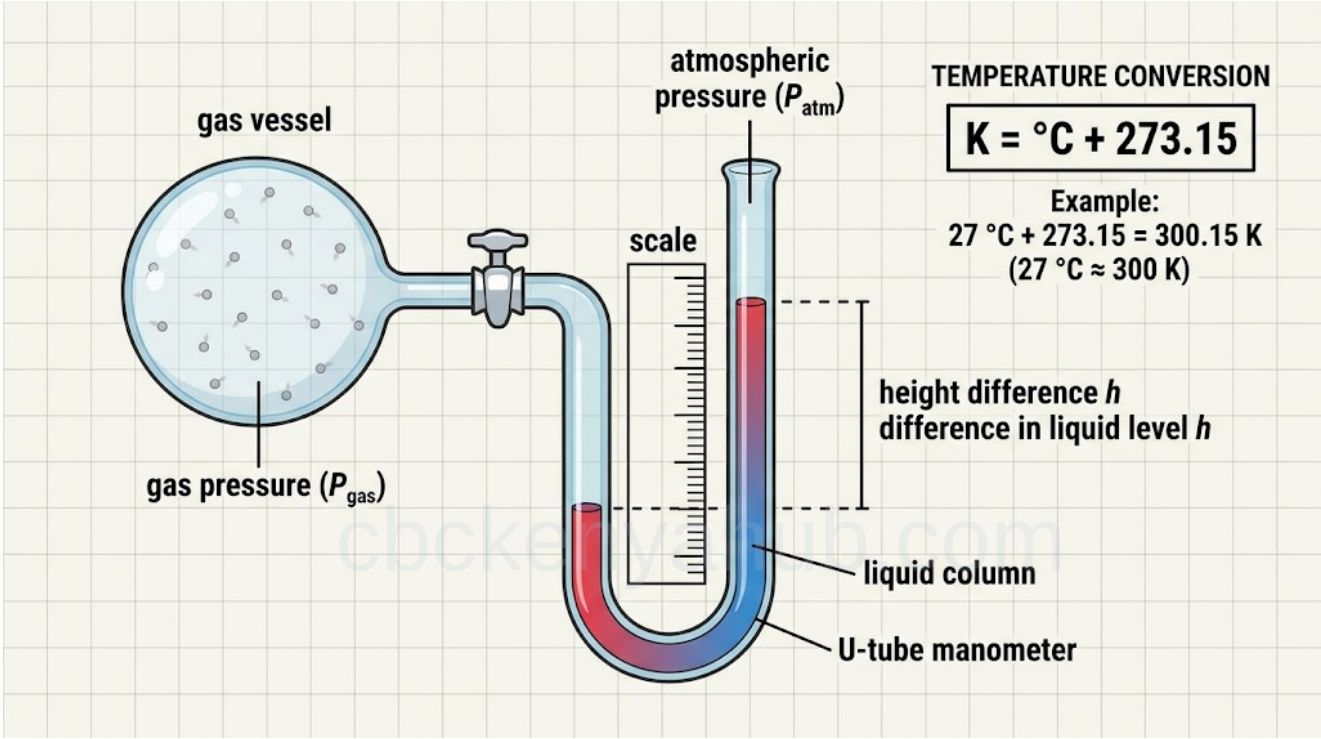
- Gas pressure is force per unit area.
- At the same temperature, ideal gas particles have the same average translational kinetic energy.
- Real gases deviate from ideal behaviour at high pressure and low temperature.



2. Pressure and Temperature

Gas pressure is measured in Pa or kPa; atmospheric pressure is about 101.3 kPa. Gas-law calculations require absolute temperature: T(K) = t(°C) + 273.15. Use consistent units and identify which variables remain constant.

- 1 atm ≈ 101.3 kPa.
- 1 dm^3 = 1 L; 1 $dm^3 = 10^{-3}$ m^3.
- Never heat a sealed container unless designed and approved for that purpose.



3. Boyle's Law

For a fixed amount of gas at constant temperature, pressure is inversely proportional to volume. When volume decreases, wall collisions become more frequent and pressure increases. Equation: $P_1V_1 = P_2V_2$.

- Use identical pressure units and identical volume units for both states.
- Temperature and amount of gas must remain constant.
- Applications include a syringe with its nozzle blocked.

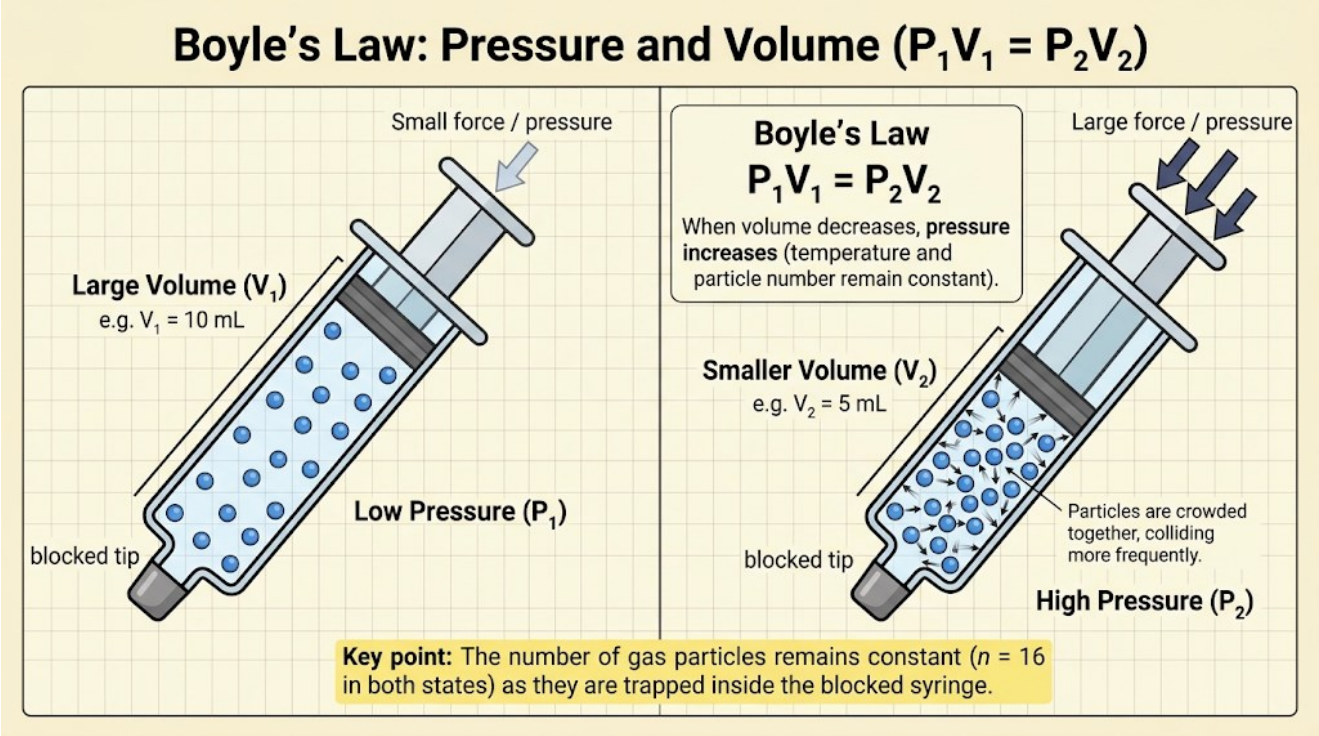
Worked Example 11: Boyle's law

Problem. A gas occupies 600 cm^3 at 100 kPa. Find volume at 150 kPa at constant temperature.

Step 1: Use $P_1V_1 = P_2V_2$.

Step 2: $V_2 = (100 \times 600)/150 = 400\ cm^3$.

Answer. The volume is 400 cm^3.



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4. Charles's Law

At constant pressure, the volume of a fixed amount of gas is directly proportional to absolute temperature: $V_1/T_1 = V_2/T_2$. Heating increases particle kinetic energy; at constant pressure the gas expands.

- Use kelvin, not Celsius.
- Doubling absolute temperature doubles volume for an ideal gas at constant pressure.
- Hot-air balloon lift also depends on the mass of the balloon and payload.

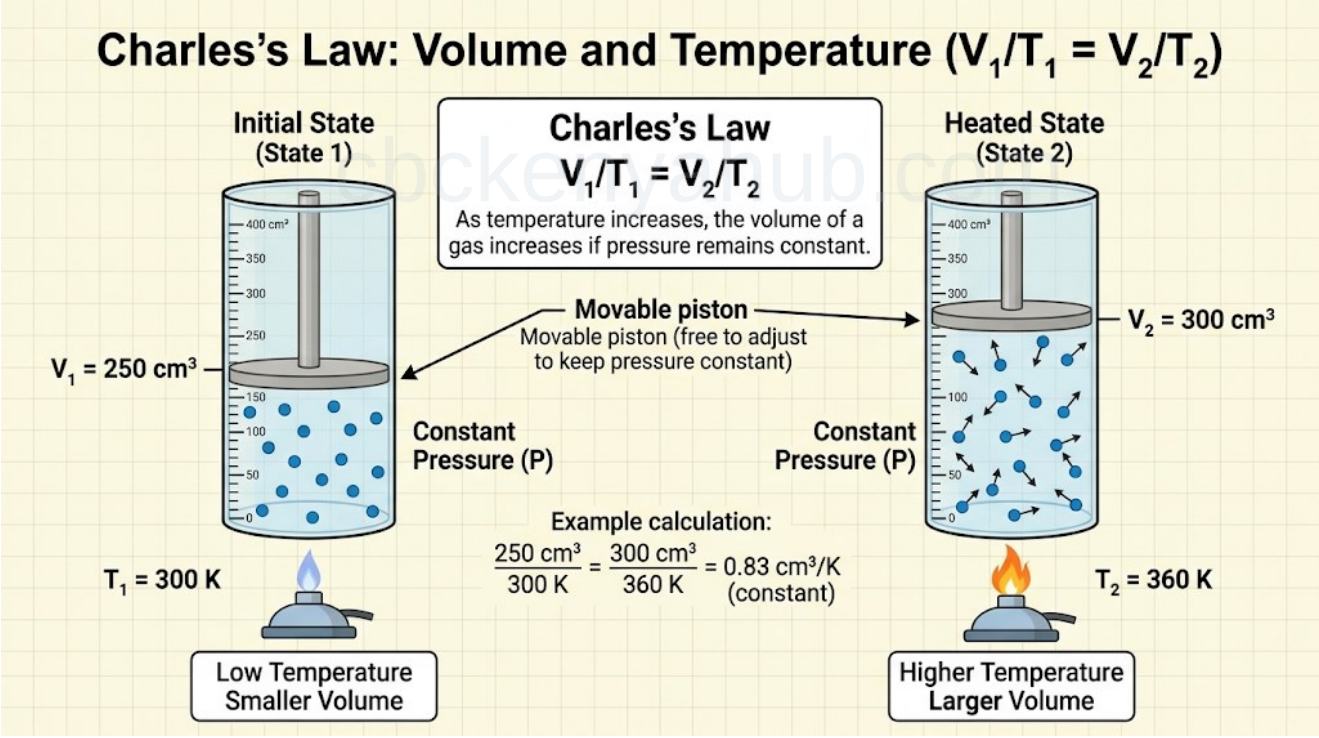
Worked Example 12: Charles's law

Problem. A gas occupies 250 cm³ at 300 K. Find its volume at 360 K at constant pressure.

Step 1: Use $V_1/T_1 = V_2/T_2$.

Step 2: $V_2 = 250 \times 360/300 = 300$ cm³.

Answer. The volume is 300 cm³.



5. Pressure Law

At constant volume, pressure is directly proportional to absolute temperature: $P_1/T_1 = P_2/T_2$. Heating a gas in a rigid container increases the frequency and force of collisions with the walls.

- The container must be rigid.

- Tyre pressure changes with temperature, though real tyres do not perfectly satisfy the simple model.
- Heating sealed containers can be dangerous.

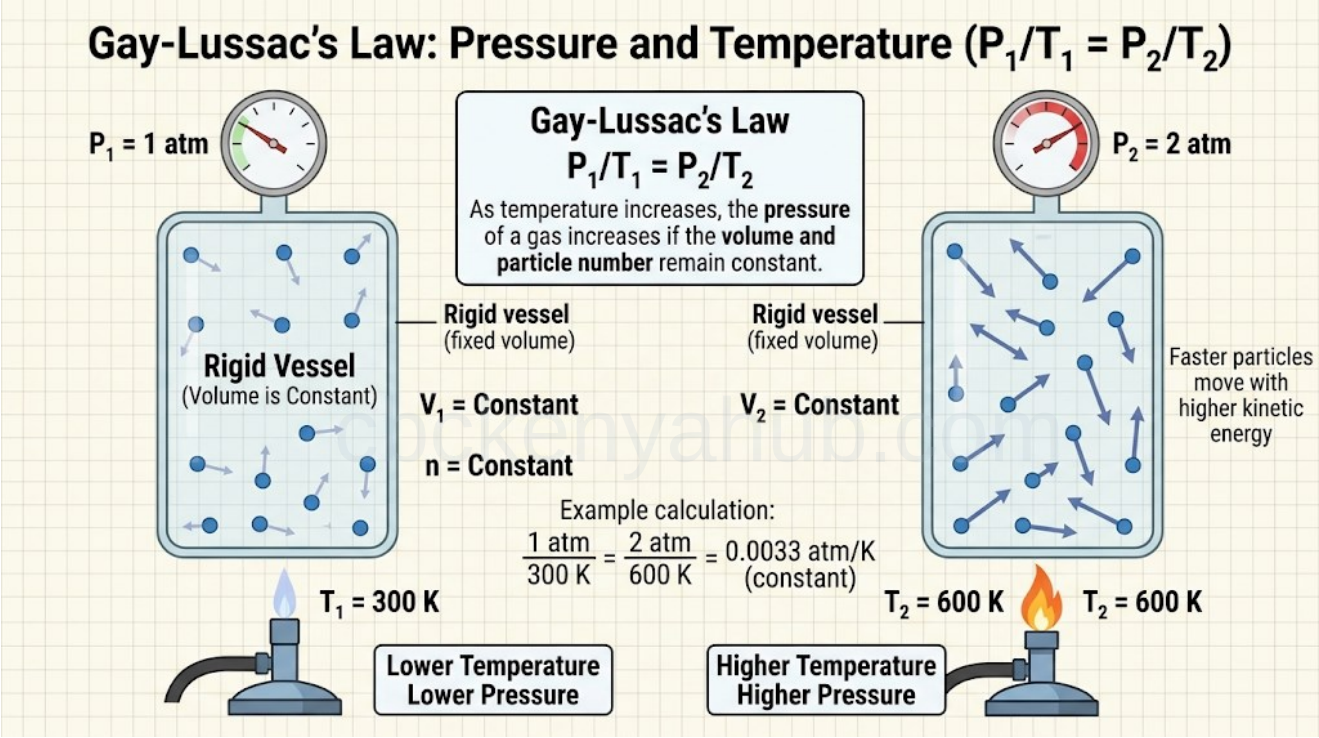
Worked Example 13: Pressure law

Problem. A gas has pressure 100 kPa at 300 K. Find pressure at 330 K in a rigid container.

Step 1: Use $P_1/T_1 = P_2/T_2$.

Step 2: $P_2 = 100 \times 330/300 = 110$ kPa.

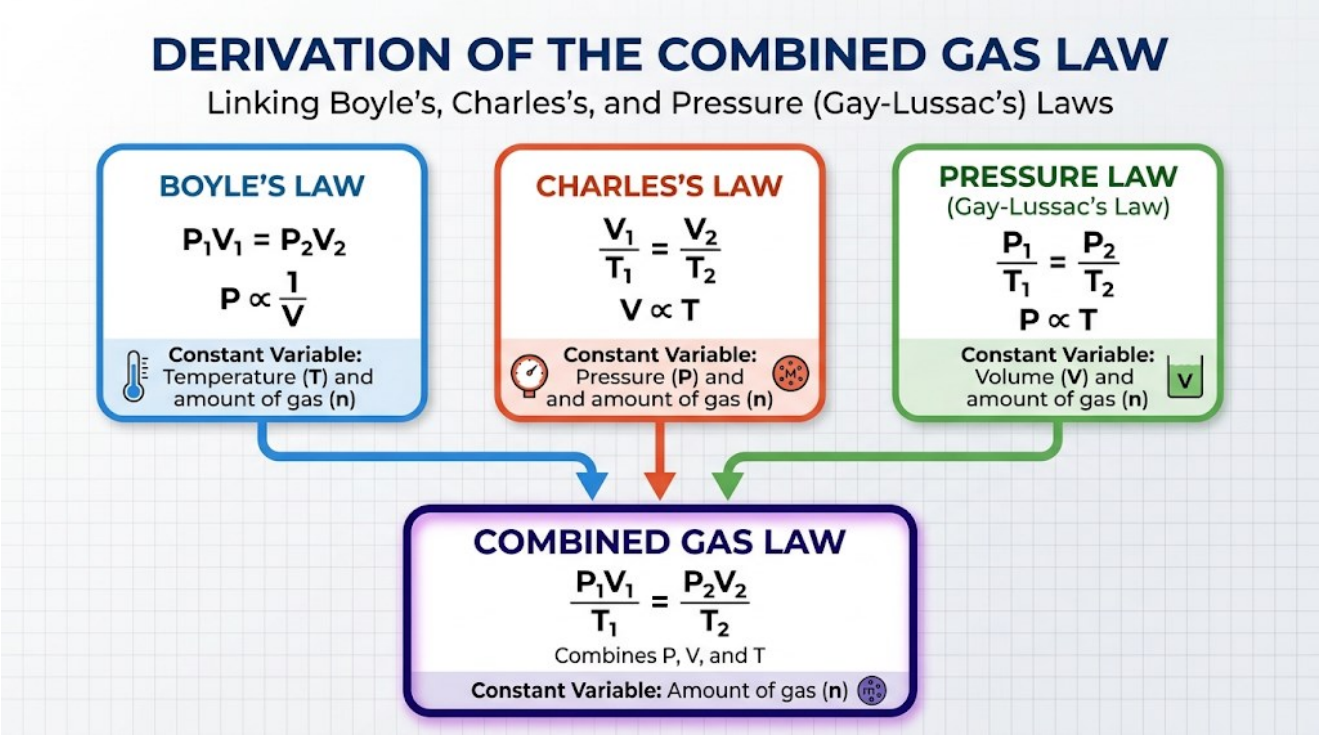
Answer. The pressure is 110 kPa.



6. Combined Gas Law

The combined gas law relates pressure, volume and temperature for a fixed amount of gas: $P_1V_1/T_1 = P_2V_2/T_2$. It combines Boyle's, Charles's and pressure laws.

- Convert temperatures to kelvin.
- Use consistent units across both states.
- If one variable is constant, the expression reduces to an individual gas law.



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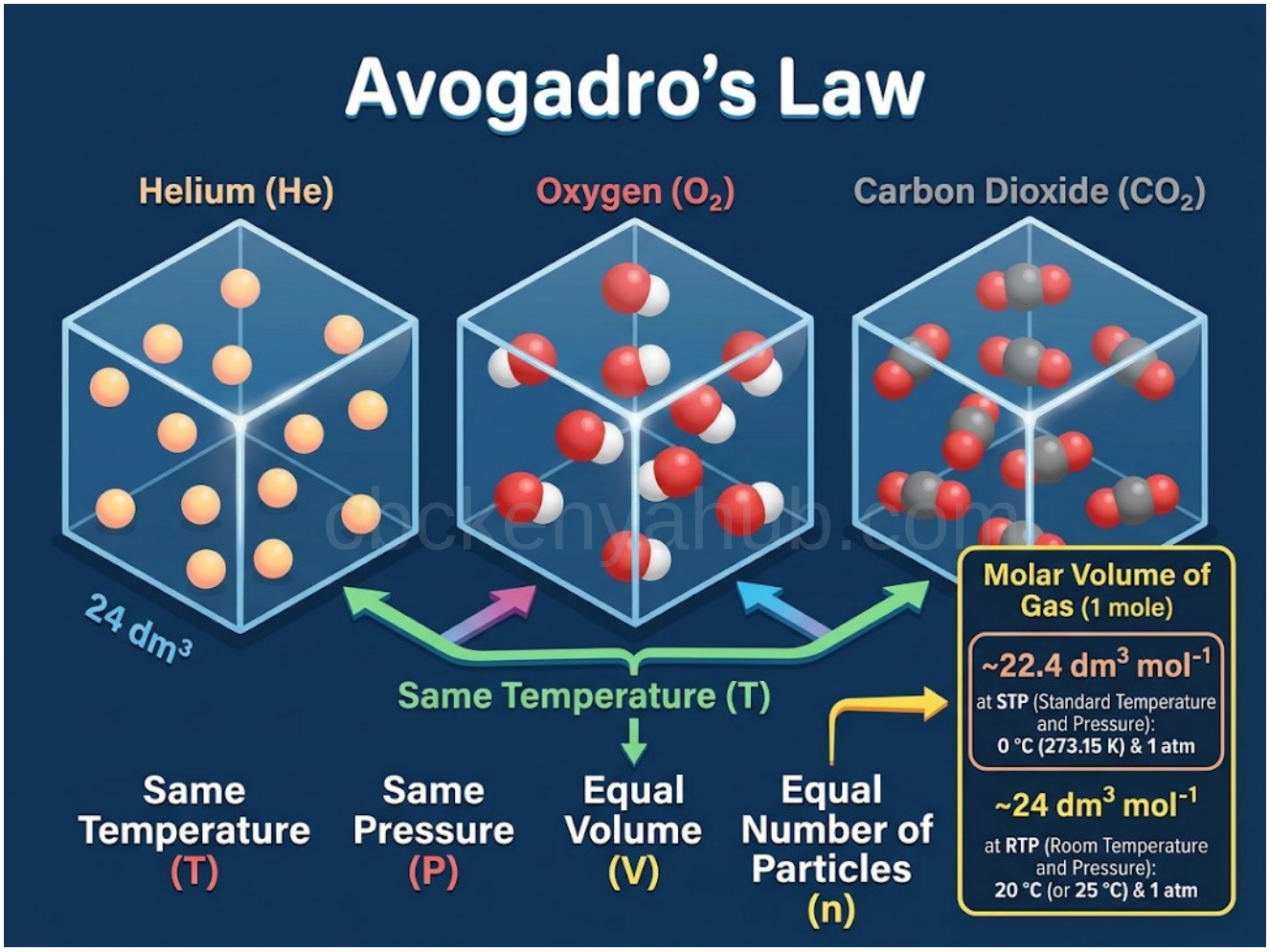
7. Avogadro's Law and Molar Volume

At fixed temperature and pressure, gas volume is directly proportional to amount in moles: V/n = constant. Equal gas volumes at the same temperature and pressure contain equal numbers of molecules.

- At room conditions, a commonly used approximate molar gas volume is 24 dm^3 mol^{-1}.
- At 0 °C and 1 atm, the ideal-gas value is about 22.4 dm^3 mol^{-1}.
- Always use the molar volume matching the stated conditions.



8. Ideal Gas Equation

The equation PV = nRT links pressure, volume, amount and absolute temperature. R must match the units used. For example, R = 8.314 kPa dm^3 mol^{-1} K^{-1} when P is kPa and V is dm^3.

- Rearrange the equation to find any unknown.
- Use n = m/M when amount is derived from mass.
- Ideal behaviour is an approximation, not a perfect description of every real gas.

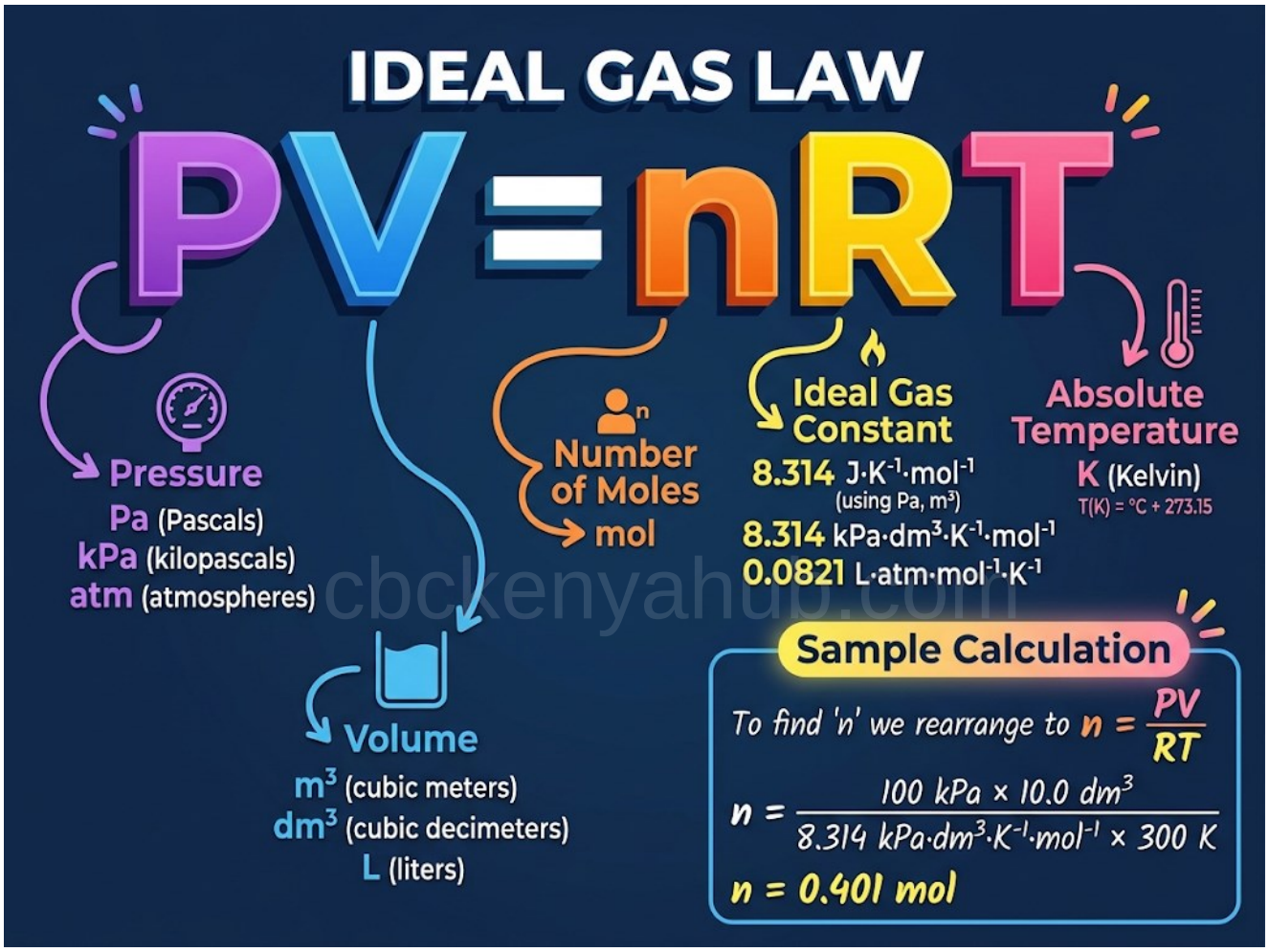
Worked Example 14: Ideal gas equation

Problem. Find moles in 10.0 dm^3 at 100 kPa and 300 K. Use R = 8.314 kPa dm^3 mol^{-1} K^{-1}.

Step 1: n = PV/RT.

Step 2: n = (100 × 10.0)/(8.314 × 300) ≈ 0.401 mol.

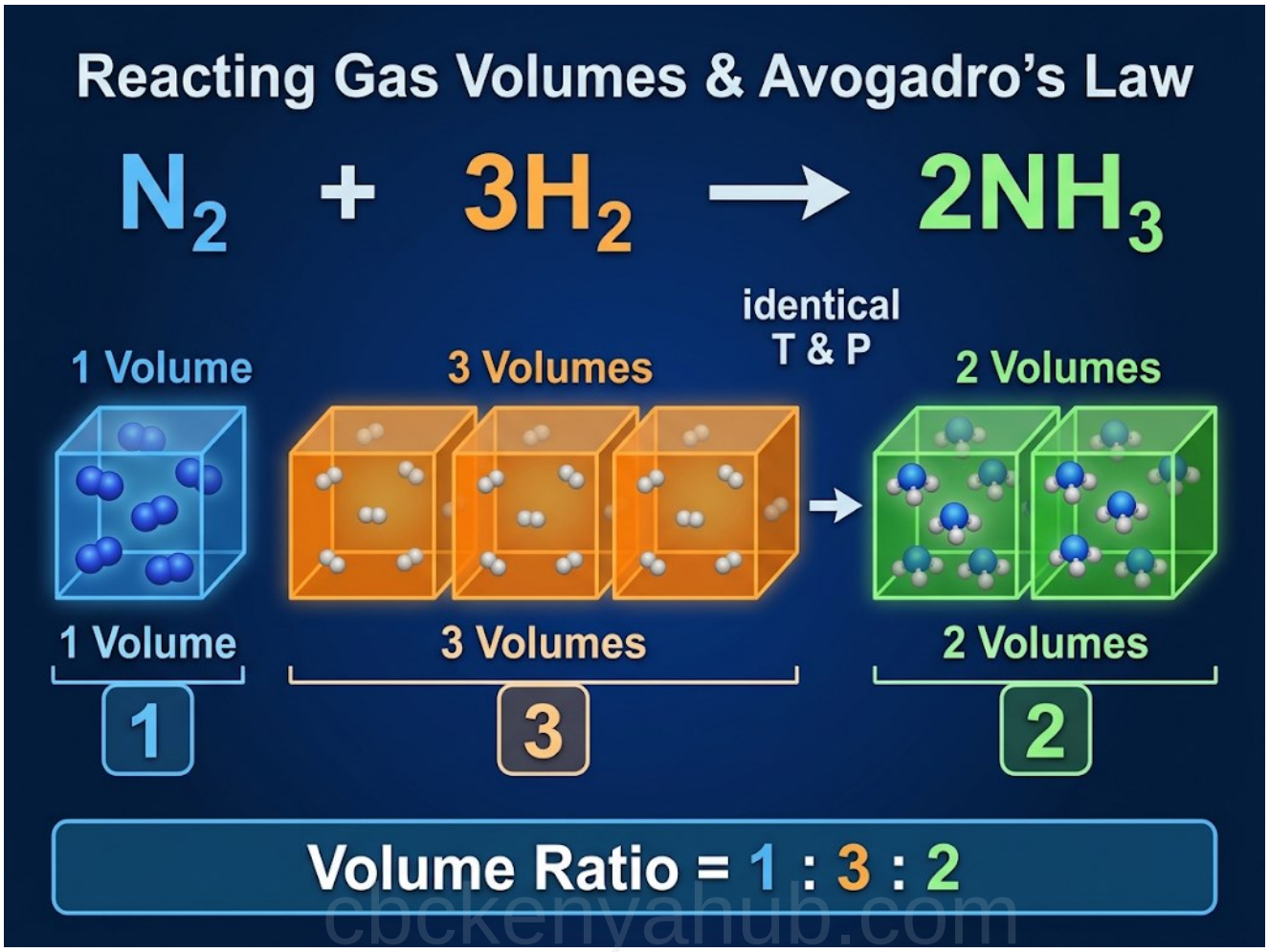
Answer. The amount is approximately 0.401 mol.



9. Gas Stoichiometry

A balanced equation gives mole ratios. At the same temperature and pressure, reacting gas volumes follow the coefficients in the equation. If conditions differ, convert volumes to moles first.

- Balance the equation.
- Use mole ratios to calculate the unknown.
- Convert the final moles to volume under stated conditions.


Reacting Gas Volumes & Avogadro's Law
N_2 + $3H_2$ → $2NH_3$
identical
T & P
1 Volume
3 Volumes
2 Volumes
1 Volume
3 Volumes
2 Volumes
1
3
2
Volume Ratio = 1 : 3 : 2
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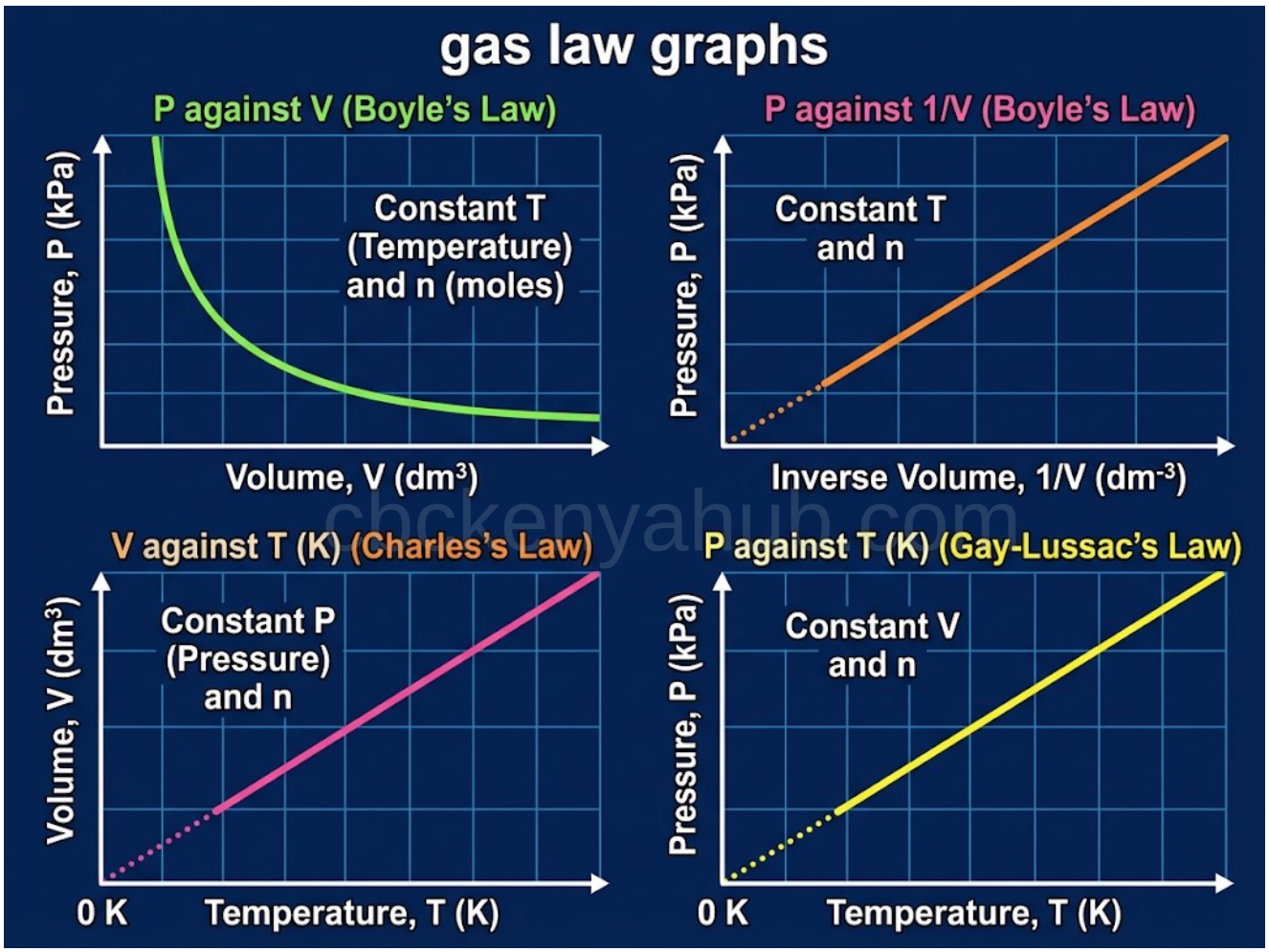
10. Graphs and Problem Solving

Boyle's law gives a curve for P against V and a straight line for P against 1/V. Charles's law gives a straight line for V against kelvin temperature; pressure against kelvin temperature is also linear at constant volume.

- Read graph axes and units.
- Convert Celsius to kelvin before applying proportionality.
- List knowns, choose the law, substitute, calculate and check units.



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Summary: Gas Laws

- Boyle: $P_1V_1 = P_2V_2$ at constant T.
- Charles: $V_1/T_1 = V_2/T_2$ at constant P.
- Pressure law: $P_1/T_1 = P_2/T_2$ at constant V.
- Combined law: $P_1V_1/T_1 = P_2V_2/T_2$.
- Avogadro: V/n is constant at fixed T and P.
- Ideal gas equation: PV = nRT; use consistent units and kelvin.

Self-check Questions: Gas Laws

- A gas occupies 500 cm³ at 120 kPa. Find its volume at 200 kPa at constant temperature.
- Convert −10 °C and 85 °C to kelvin.
- A gas occupies 3.0 dm³ at 250 K. Find volume at 400 K at constant pressure.
- Explain why pressure rises when gas is heated in a rigid container.
- State two assumptions of the ideal gas model.
- Calculate pressure of 0.25 mol gas in 5.0 dm³ at 300 K.

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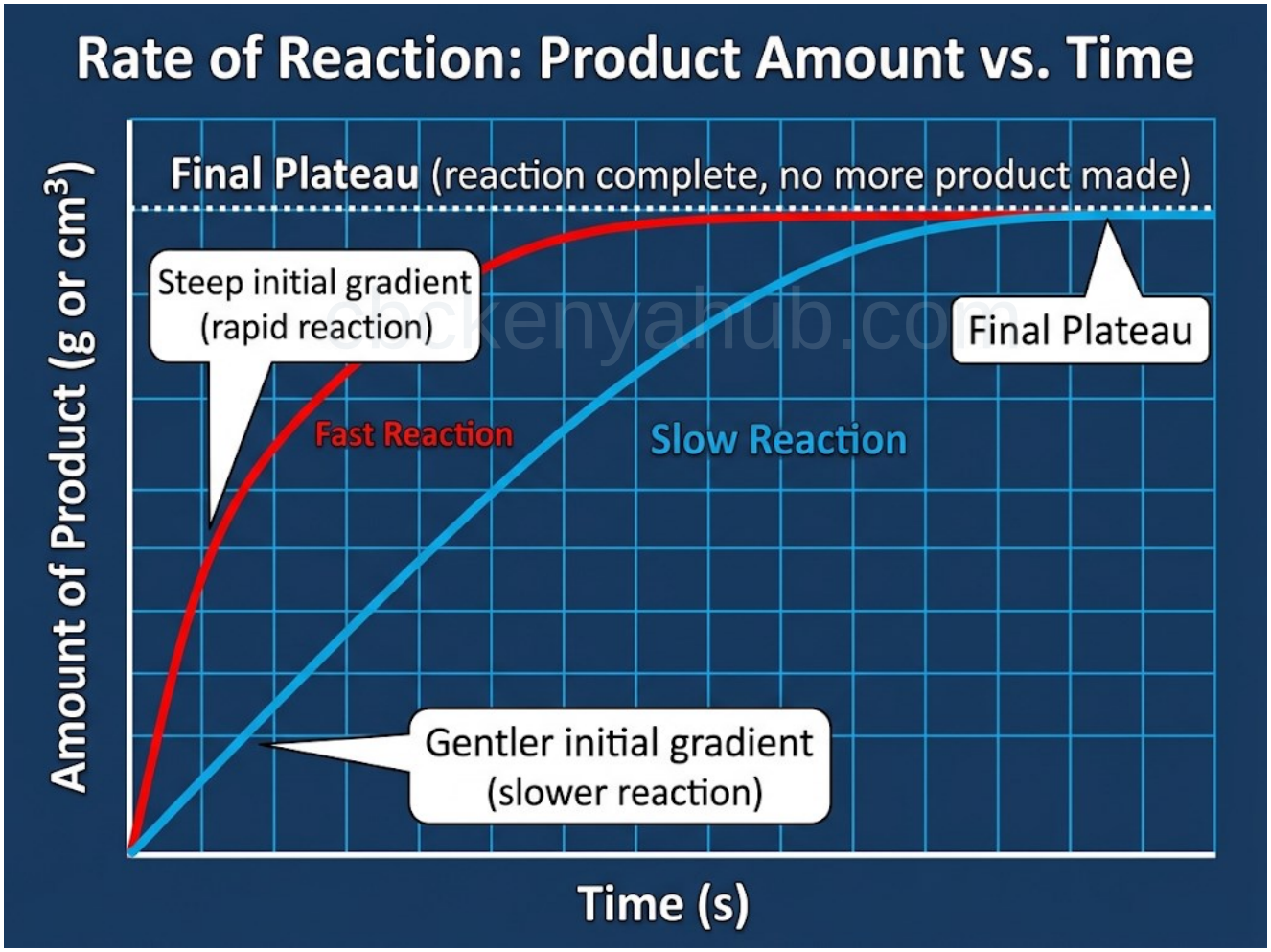
SUB-STRAND 2.3: REACTION RATES

Curriculum focus: investigate factors affecting reaction rates, explain trends using particle ideas and collision theory, and relate rate control to everyday life and industrial processes.

1. Meaning of Reaction Rate

Reaction rate describes how quickly reactants are consumed or products are formed. It may be measured by gas volume, mass change, colour intensity, concentration or time to a visible endpoint. A fair comparison requires similar starting conditions.

- Fast reactions include many precipitation reactions; rusting is usually slow.
- Average rate = change in measured quantity divided by elapsed time.
- The chosen measurement must be suitable for the reaction.



2. Measuring Gas Production

A gas-producing reaction can be followed using a gas syringe or, where appropriate, displacement of water. Alternatively, mass loss can be recorded if gas escapes. Plot the measured quantity against time to compare rates.

- Check for leaks and assemble apparatus before mixing.
- Record readings at regular intervals.
- Initial gradient estimates initial rate; units may be $cm^3\ s^{-1}$.

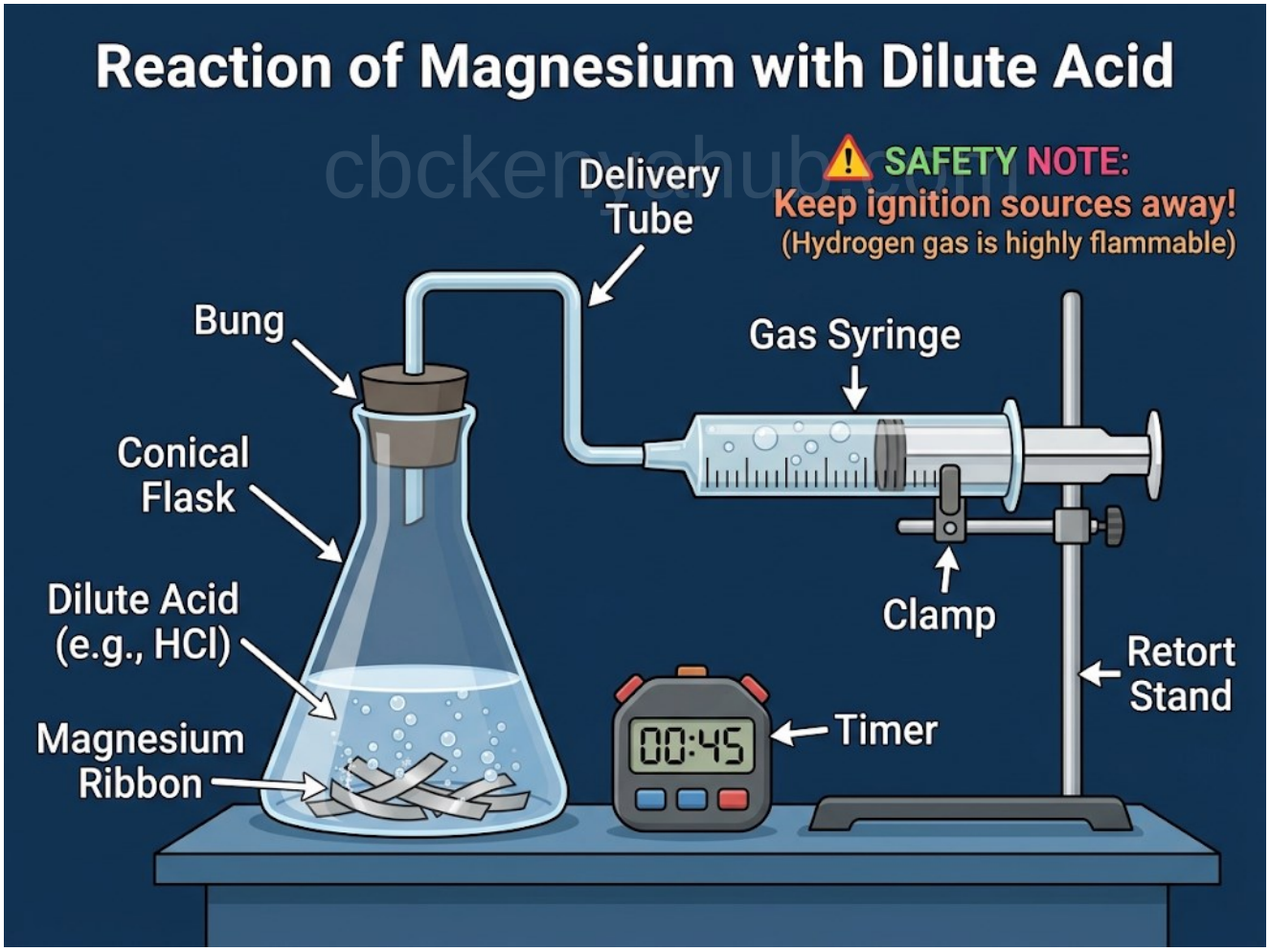
Worked Example 15: Average rate

Problem. A reaction produces 60 cm^3 gas in 30 s. Calculate the average rate.

Step 1: Rate = 60/30.

Step 2: Rate = 2.0 $cm^3\ s^{-1}$.

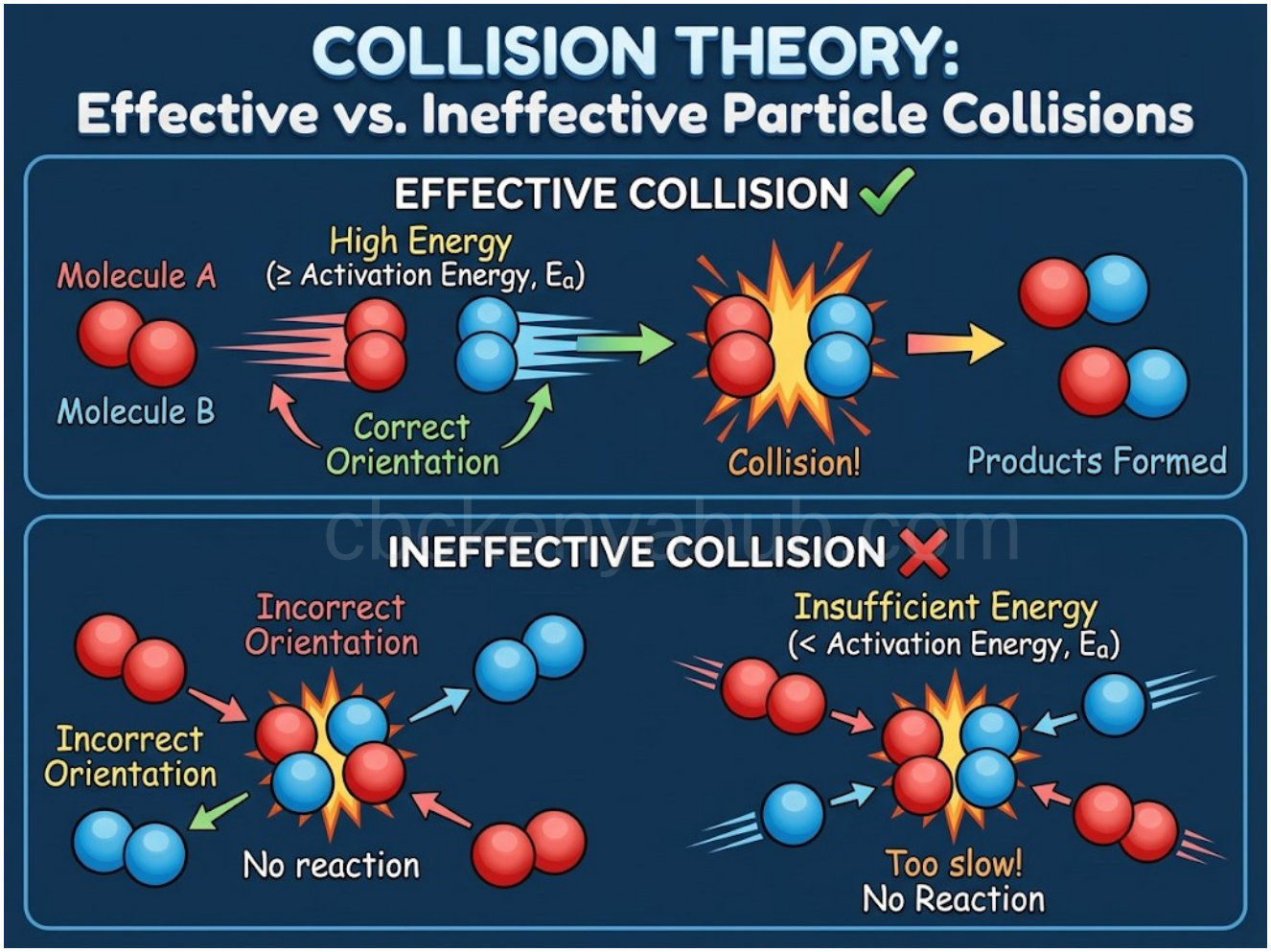
Answer. The average gas-production rate is 2.0 $cm^3\ s^{-1}$.



3. Collision Theory

Reacting particles must collide with sufficient energy and suitable orientation for a reaction to occur. Activation energy is the minimum energy required for the reaction pathway. Only effective collisions lead to product formation.

- More frequent collisions can increase rate if collisions have suitable energy and orientation.
- A catalyst provides an alternative pathway with lower activation energy.
- Temperature changes the fraction of particles with energy at or above activation energy.



4. Concentration

Increasing concentration usually increases rate because more reacting particles occupy each unit volume, causing more frequent collisions. The exact mathematical relationship depends on the reaction and is established experimentally.

- Keep temperature, solid surface area and total volume controlled where relevant.
- Use safe, teacher-approved concentrations.
- Do not assume rate is always directly proportional to concentration.

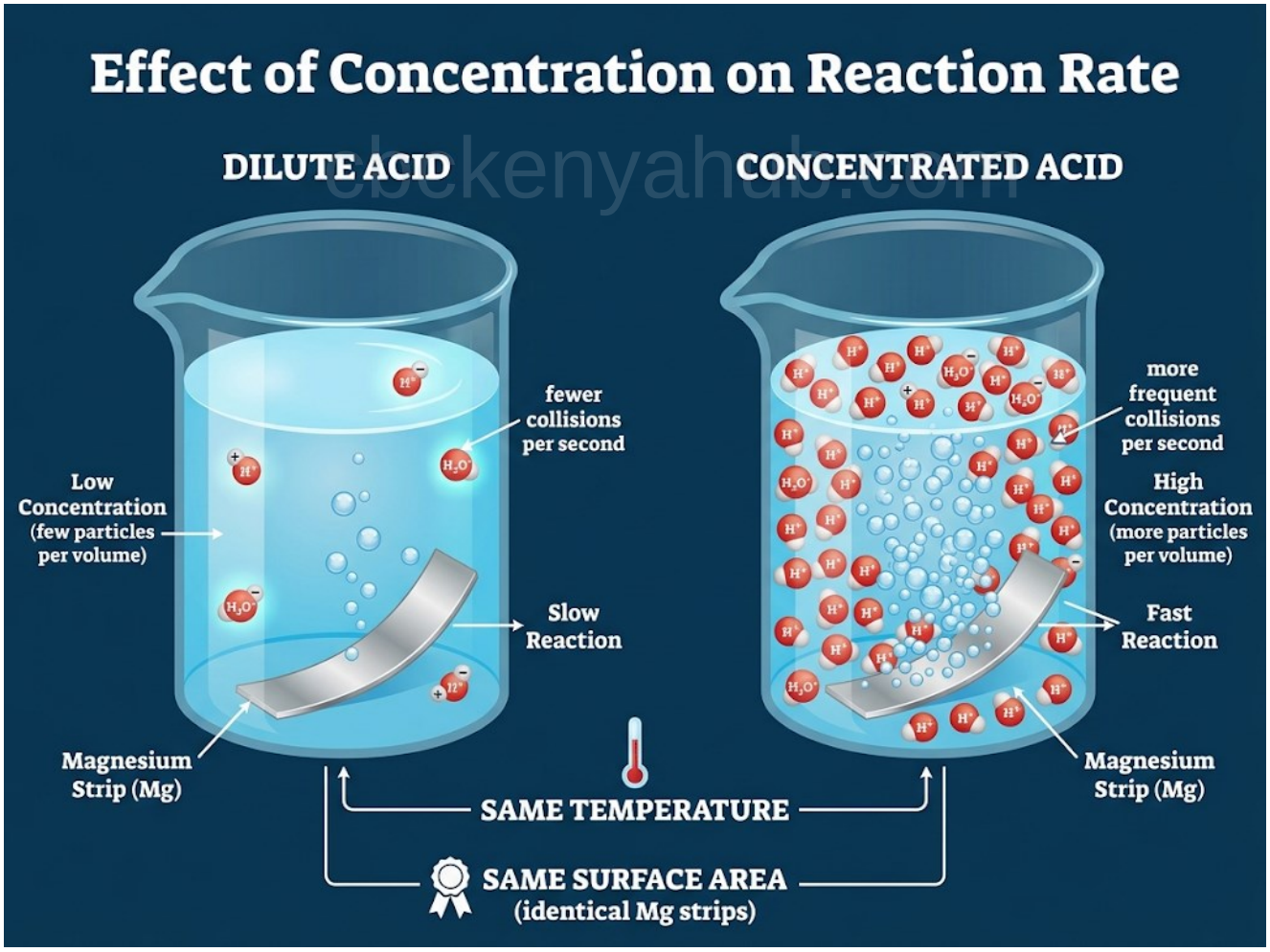
Worked Example 16: Concentration trend

Problem. Why does magnesium react faster with 2.0 mol dm^{-3} HCl than with 0.5 mol dm^{-3} HCl?

Step 1: The more concentrated solution contains more acid particles per unit volume.

Step 2: More frequent effective collisions occur at the magnesium surface.

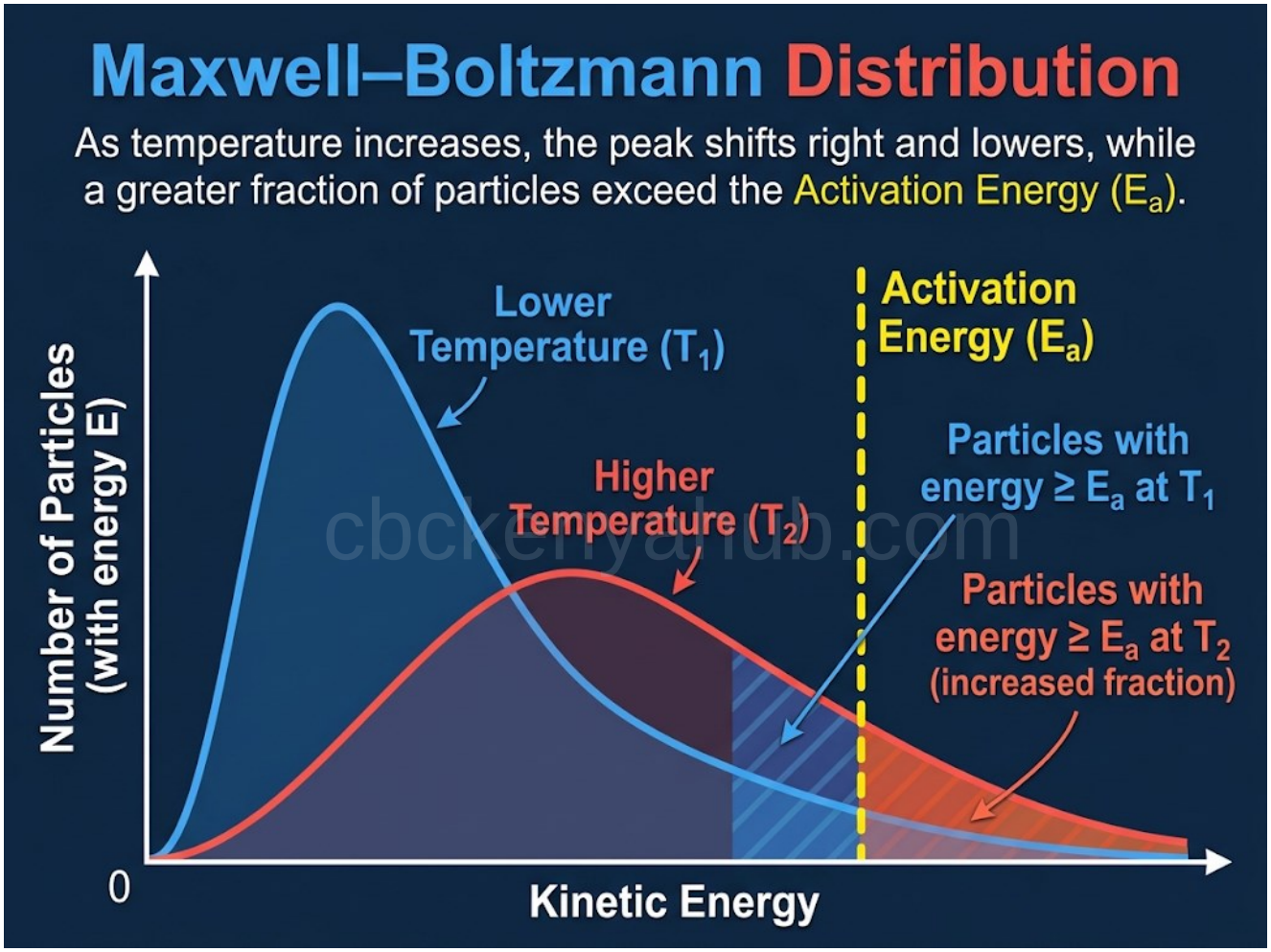
Answer. The more concentrated acid generally gives a faster reaction, all else equal.



5. Temperature

Increasing temperature usually increases reaction rate. Particles move faster and collide more frequently; more importantly, a larger fraction of particles have energy sufficient to overcome activation energy.

- Allow solutions to reach the intended temperature before mixing.
- A water bath can help maintain temperature.
- Do not claim every reaction doubles in rate for a fixed temperature rise.



6. Surface Area

Breaking a solid into smaller pieces increases its exposed surface area. More particles of the solid are available for contact with another reactant, so the rate usually increases.

- Compare equal masses and keep concentration, volume and temperature constant.
- This factor is most relevant when a solid reacts with a liquid or gas.
- Fine powders may present dust and rapid-reaction hazards.

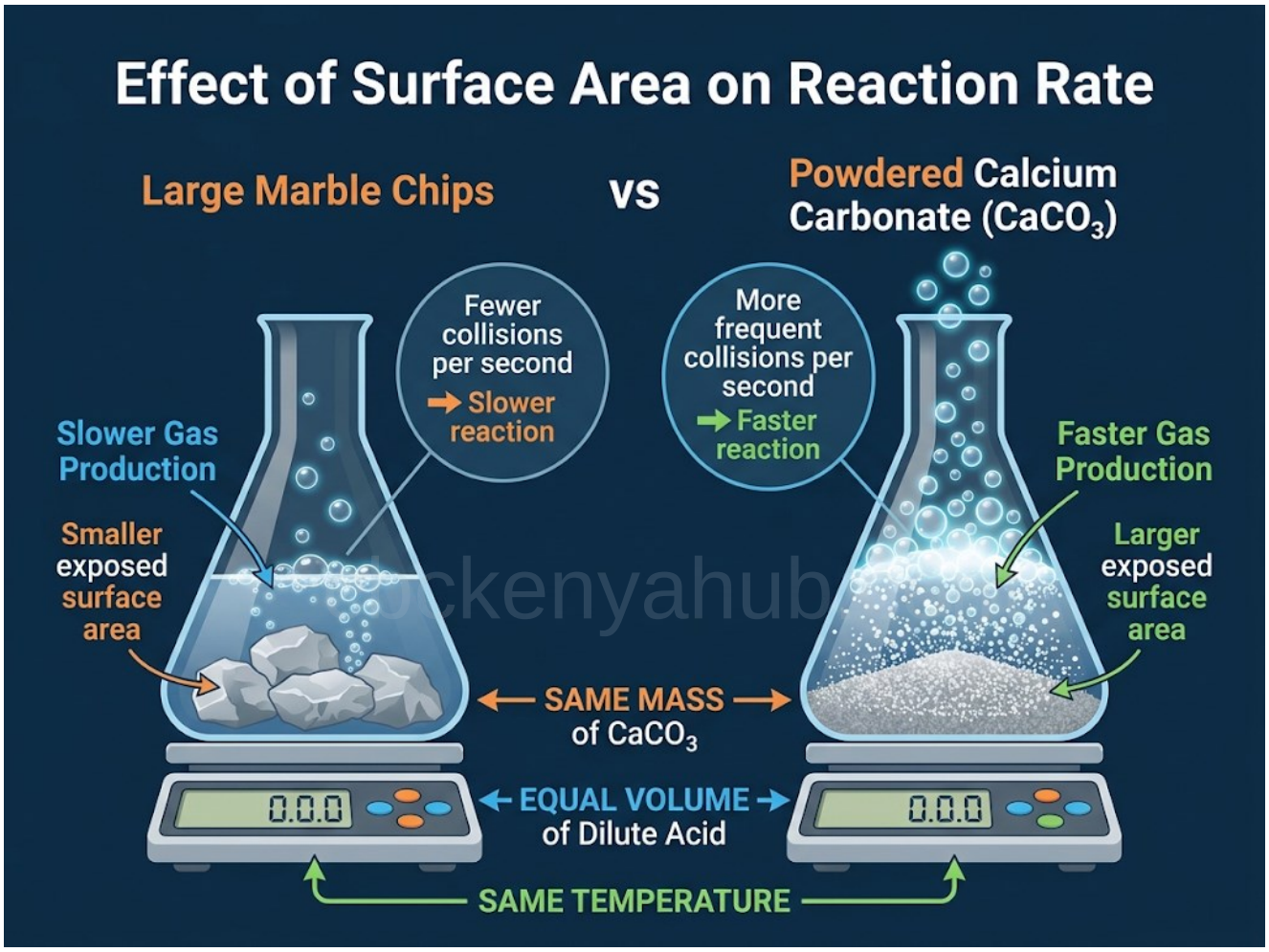
Worked Example 17: Surface area

Problem. Why does powdered calcium carbonate react faster than large chips of the same mass?

Step 1: Powder exposes a larger total surface area.

Step 2: More acid particles can contact the solid at the same time.

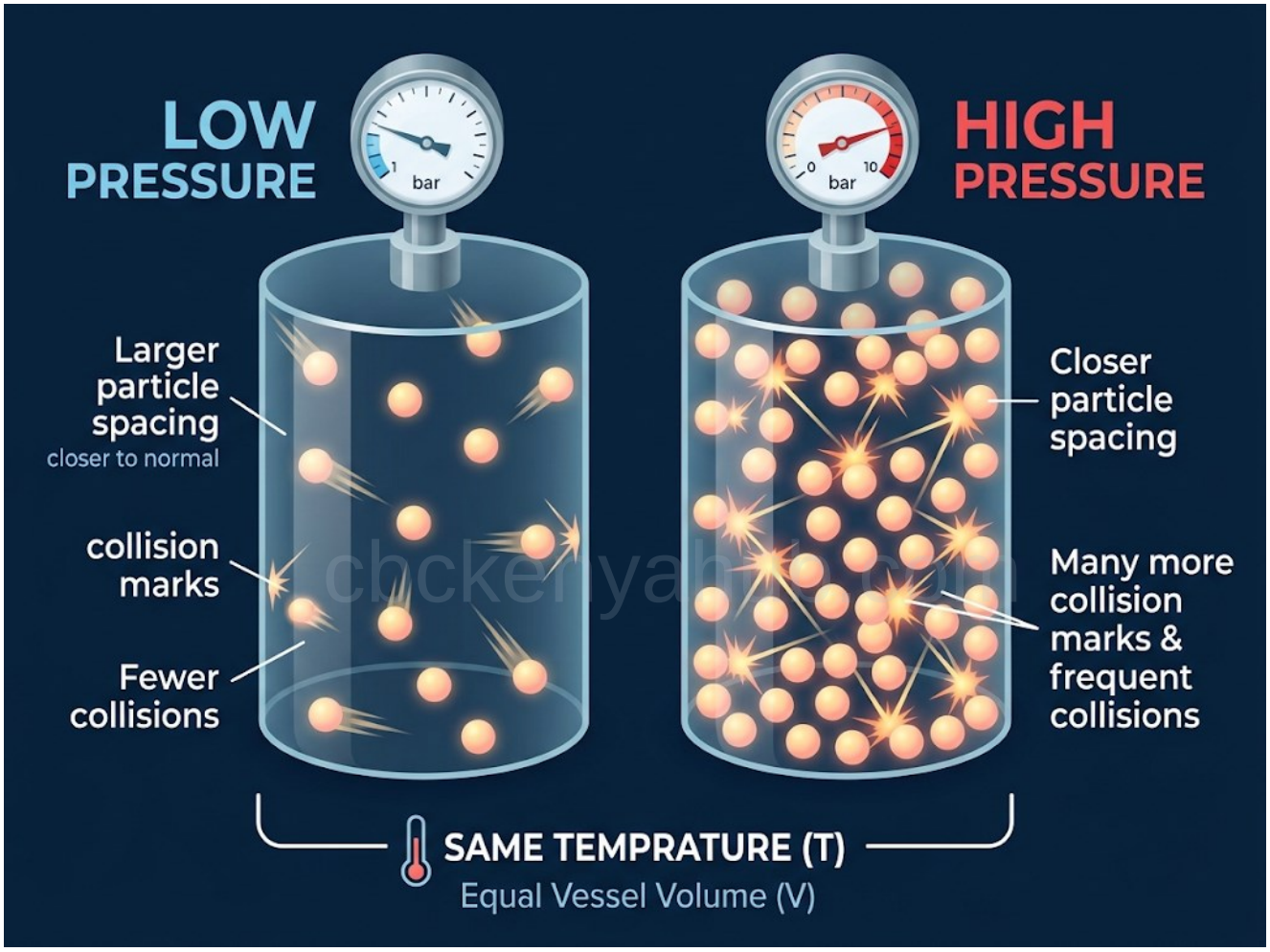
Answer. The reaction is faster because collision frequency at the interface is greater.



7. Pressure for Gaseous Reactants

Increasing pressure at constant temperature generally increases gas concentration per unit volume. The particles are closer together and collide more frequently, so gas-phase reaction rate may increase.

- The effect is most relevant to gaseous reactants.
- Pressure has little direct effect on reactions involving only solids and liquids.
- Faster rate does not necessarily mean greater equilibrium yield.



8. Catalysts and Activation Energy

A catalyst increases reaction rate by providing an alternative pathway with lower activation energy. It is regenerated overall and does not change the overall energy difference between reactants and products or the equilibrium position.

- Examples include manganese(IV) oxide in hydrogen peroxide decomposition, iron in the Haber process and catalytic converters.
- Enzymes are biological catalysts.
- Catalysts may lose activity or be poisoned by impurities.

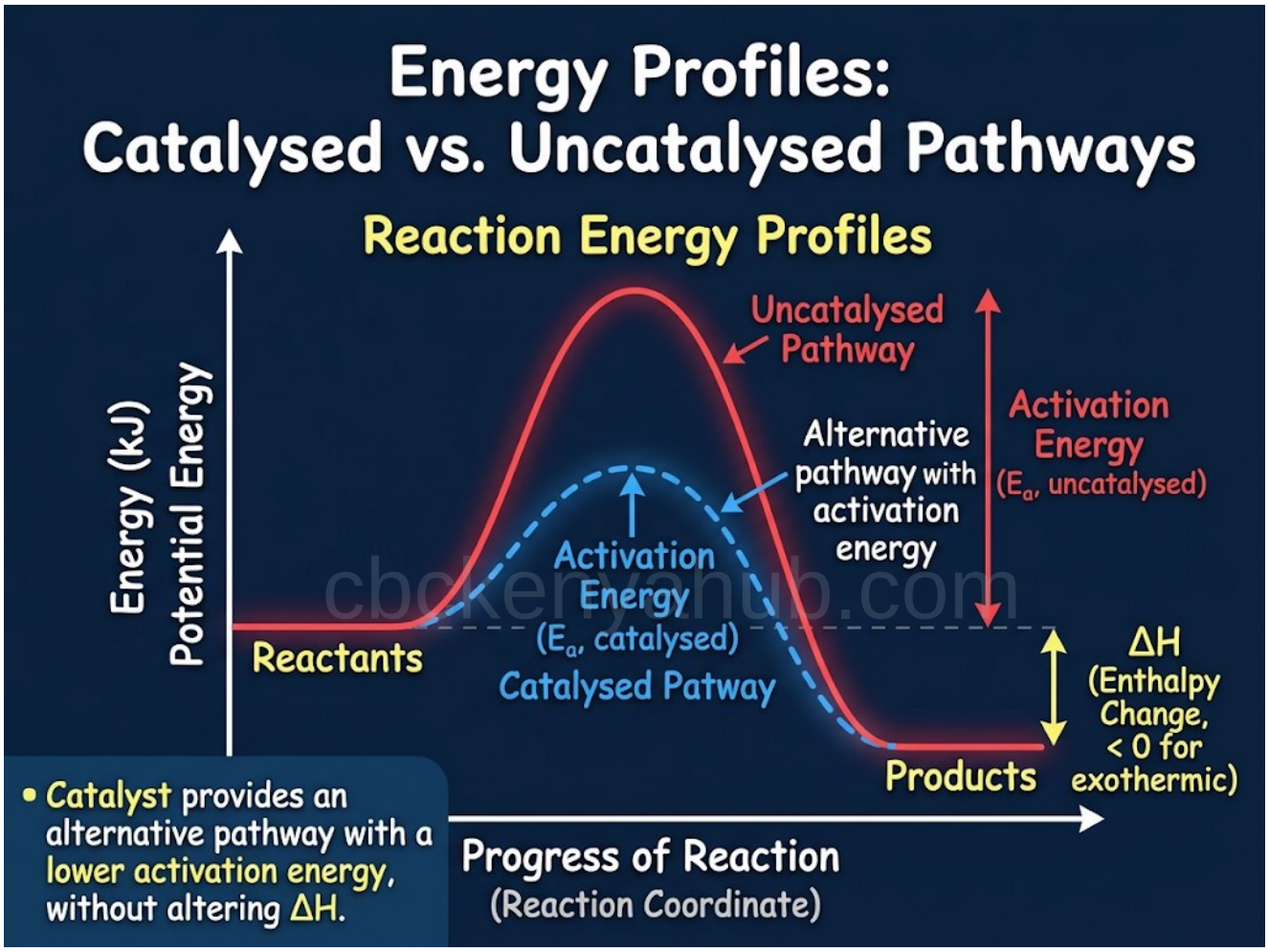
Worked Example 18: Catalyst

Problem. Explain the role of manganese(IV) oxide in hydrogen peroxide decomposition.

Step 1: It provides an alternative pathway with lower activation energy.

Step 2: The observed bubbles are oxygen gas; the catalyst is regenerated overall.

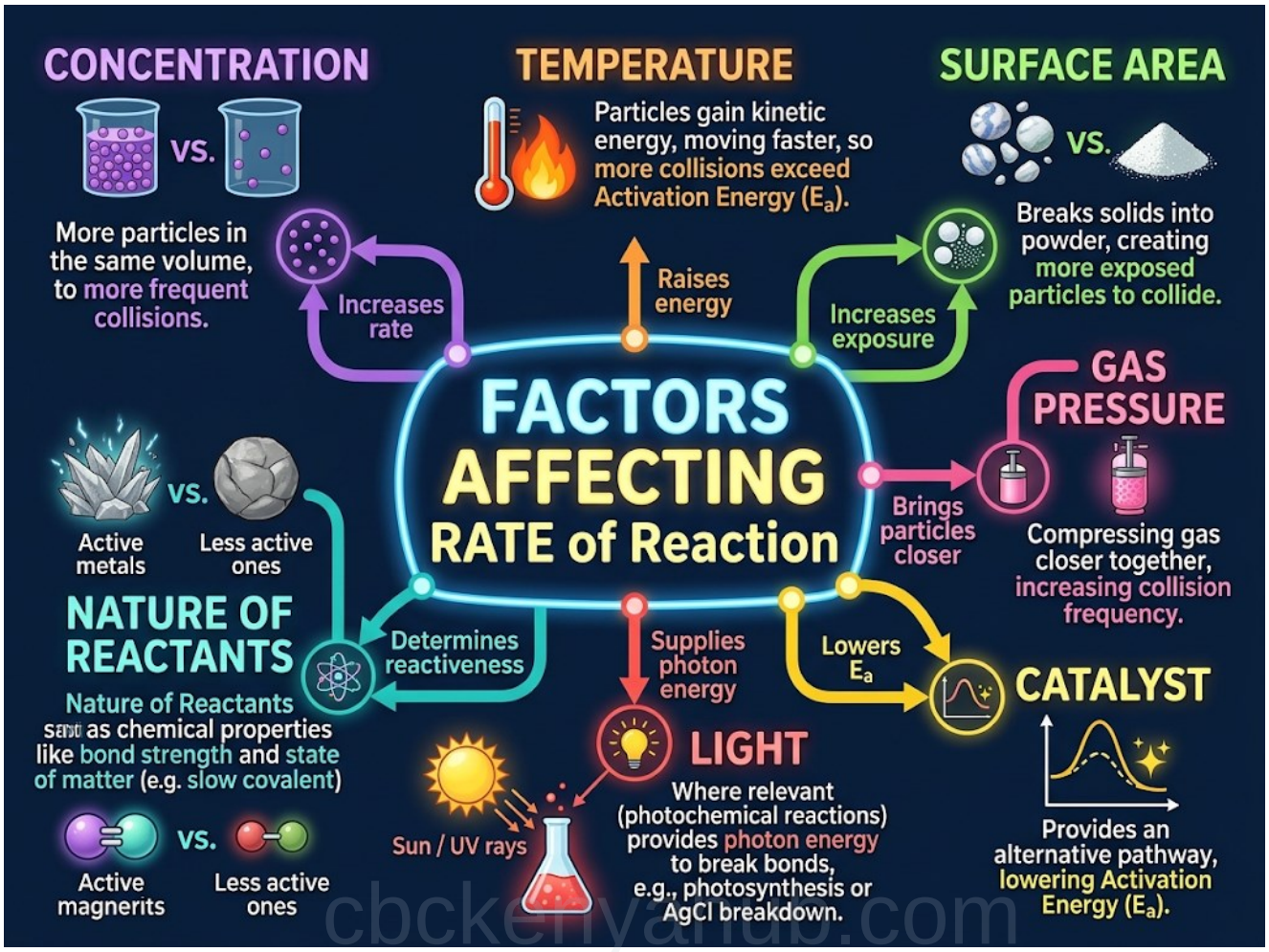
Answer. The reaction rate increases but the overall products do not change.



9. Nature of Reactants and Other Factors

Rate also depends on chemical nature, physical state and mechanism. Ionic reactions in solution can be rapid; reactions requiring bond breaking may be slower. Light can initiate some photochemical reactions.

- Different metals may react at different rates with the same acid.
- Stirring can improve contact between a solid and solution.
- Use clean apparatus and consistent reagents.



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10. Rate Graphs

A product-against-time graph generally rises and then levels off; a reactant-against-time graph generally falls. The gradient represents rate. A steep gradient indicates a fast rate and a plateau indicates little further measurable change.

- Average rate = change in quantity/time.
- Use a tangent to estimate instantaneous rate where required.
- Include units such as $cm^3\ s^{-1}$, $g\ s^{-1}$ or $mol\ dm^{-3}\ s^{-1}$.

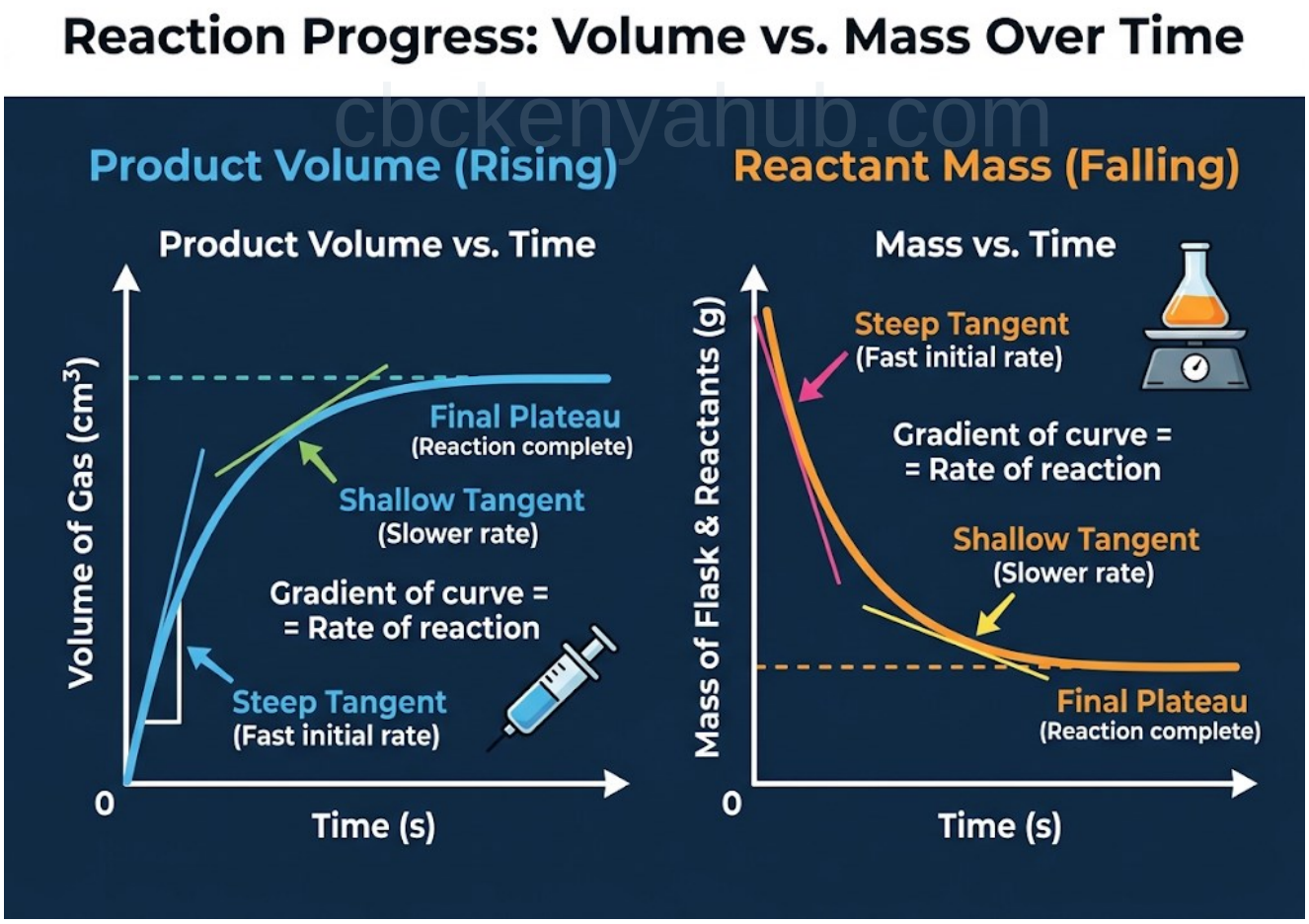
Worked Example 19: Average rate from readings

Problem. Gas volume rises from 12 cm^3 at 10 s to 36 cm^3 at 20 s. Find average rate.

Step 1: Change in volume = 36 − 12 = 24 cm^3.

Step 2: Elapsed time = 10 s; rate = 24/10 = 2.4 $cm^3\ s^{-1}$.

Answer. Average rate is 2.4 $cm^3\ s^{-1}$.



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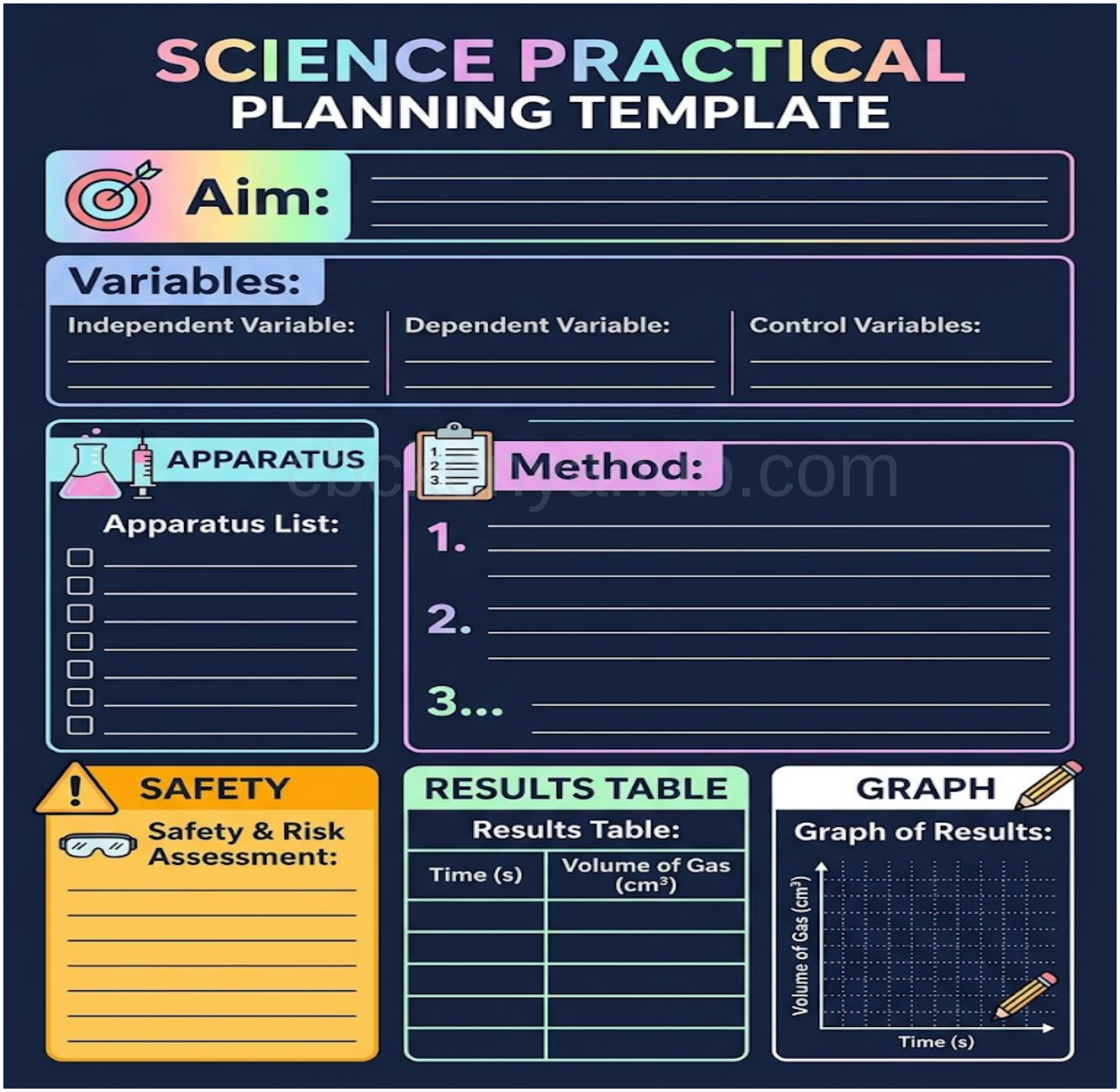
11. Fair Investigations

A fair test changes one independent variable, measures a dependent variable and controls other relevant factors. Repeat measurements to improve reliability and identify anomalous results.

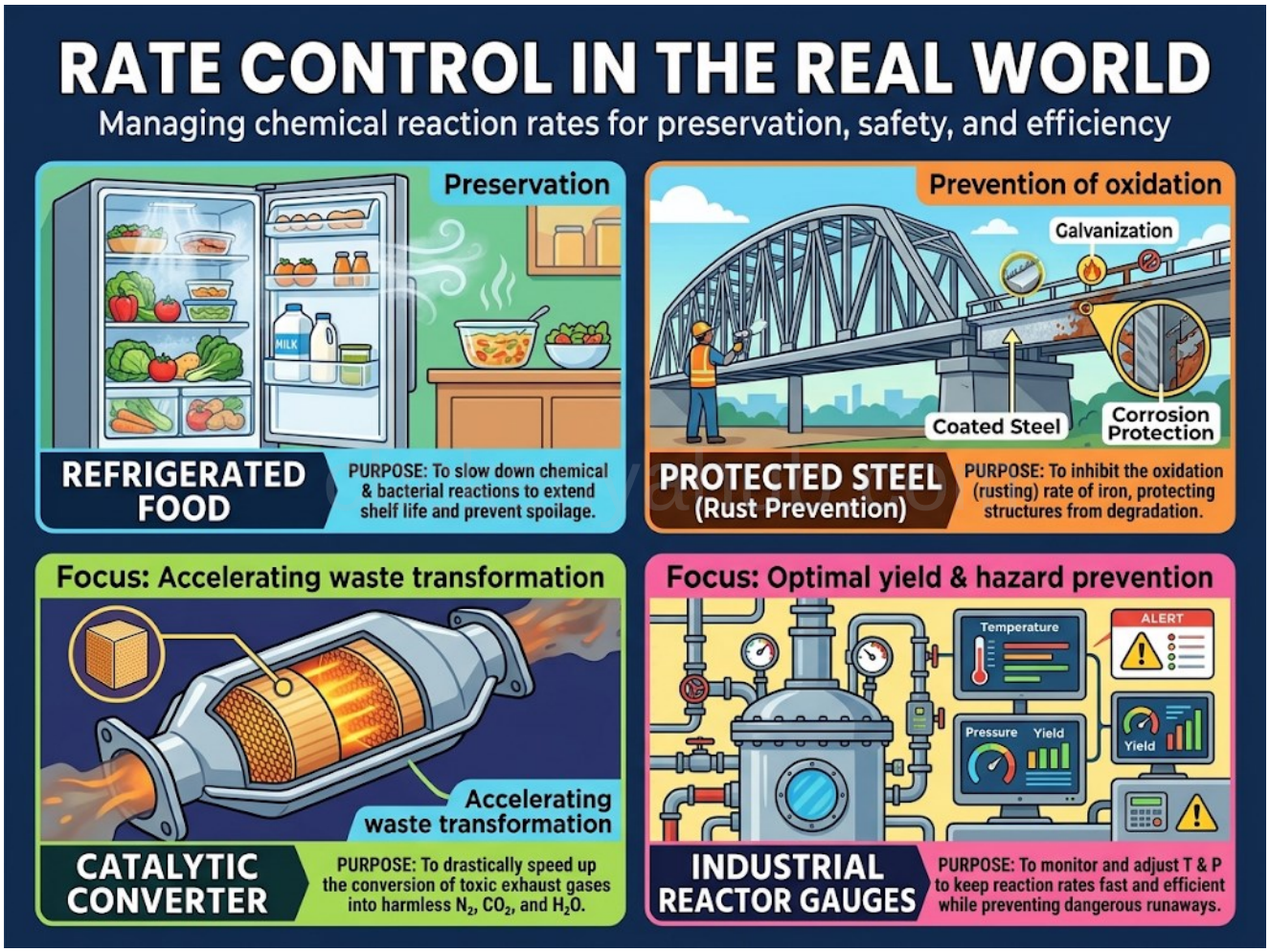
- Independent variable: factor deliberately changed.
- Dependent variable: measured outcome.
- Controlled variables: conditions held constant.
- Record results immediately and use a suitable graph.



12. Importance of Rate Control

Rate control matters in manufacturing, food preservation, medicine, environmental management and safety. Industry may increase rates for productivity, while refrigeration and protective coatings slow unwanted processes.

- Refrigeration slows many food-spoilage processes.
- Paint or galvanising can reduce rusting.
- Catalytic converters speed up reactions that reduce harmful exhaust components.
- Safe operation balances rate, product quality, cost and risk.



Summary: Reaction Rates

- Rate is the change in a measurable reactant or product quantity per unit time.
- Effective collisions need sufficient energy and suitable orientation.
- Higher concentration, temperature, surface area and gas pressure often increase rate under suitable conditions.
- Catalysts lower activation energy and are regenerated overall.

- Graphs show rate through their gradient.
- Fair tests change one variable and control other relevant conditions.

Self-check Questions: Reaction Rates

- Define reaction rate and name two measurable quantities.
- Explain the temperature effect using collision theory.
- Why does powdered calcium carbonate react faster than large chips?
- Explain a catalyst using activation energy.
- A reaction produces 90 cm^3 gas in 45 s. Calculate average rate.
- Plan a fair test of concentration and reaction rate.

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STRAND 2.0: CONSOLIDATION AND REVISION

Key Formulae at a Glance

Concept	Relationship	Conditions / reminders
Boyle's law	$P_1V_1 = P_2V_2$	Fixed amount; constant T
Charles's law	$V_1/T_1 = V_2/T_2$	Fixed amount; constant P; T in K
Pressure law	$P_1/T_1 = P_2/T_2$	Fixed amount; constant V; T in K
Combined gas law	$P_1V_1/T_1 = P_2V_2/T_2$	Fixed amount; T in K
Ideal gas equation	$PV = nRT$	Consistent units; T in K
Average rate	Change in quantity ÷ time	Include units
Moles from molar volume	$n = V/V_m$	Use V_m at stated conditions

Common Errors to Avoid

- Writing salt formulae without balancing ionic charges.
- Assuming every salt is soluble or every salt solution has pH 7.
- Forgetting to wash precipitates or overheating crystals.
- Splitting a solid precipitate into ions in a net ionic equation.
- Using Celsius instead of kelvin in gas-law calculations.
- Mixing units in PV = nRT without checking R.
- Claiming a catalyst changes equilibrium position.
- Ignoring activation energy when explaining the temperature effect.
- Changing several variables in one reaction-rate experiment.

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Worked Example 20: Gas law application

Problem. A 4.0 dm^3 gas sample at 100 kPa and 300 K is heated to 450 K at constant pressure. Find its new volume.

Step 1: Use Charles's law: $V_1/T_1 = V_2/T_2$.

Step 2: $V_2 = 4.0 \times 450/300 = 6.0\ dm^3$.

Answer. The new volume is 6.0 dm^3.

Worked Example 21: Precipitation

Problem. Aqueous lead(II) nitrate is mixed with aqueous potassium iodide. Predict the precipitate and net ionic equation.

Step 1: Lead(II) iodide and potassium nitrate are possible products.

Step 2: Nitrates and potassium salts are soluble; PbI_2 is insoluble and yellow.

Step 3: Spectator ions are K^+ and NO_3^-.

Answer. $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$; net ionic: $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$.

Worked Example 22: Rate comparison

Problem. Gas volume increases from 0 to 18 cm³ in the first 6 s, then from 18 to 30 cm³ in the next 6 s. Compare average rates.

Step 1: First interval: 18/6 = 3.0 $cm^3\ s^{-1}$.

Step 2: Second interval: (30 − 18)/6 = 2.0 $cm^3\ s^{-1}$.

Answer. The rate decreased from 3.0 to 2.0 $cm^3\ s^{-1}$.

End-of-Strand Revision Test

- 1. Define a salt and distinguish between normal and acid salts.
- 2. Write formulae for magnesium nitrate, iron(III) sulfate and ammonium carbonate.
- 3. Describe how to prepare copper(II) sulfate crystals from copper(II) oxide and dilute sulfuric acid.
- 4. Explain precipitation and why a precipitate is washed.
- 5. Write the net ionic equation for $Ba^{2+}(aq)$ reacting with $SO_4^{2-}(aq)$.
- 6. A gas occupies 750 cm³ at 100 kPa. Find its volume at 125 kPa at constant temperature.
- 7. A gas occupies 2.4 dm³ at 280 K. Find its volume at 350 K at constant pressure.
- 8. State two assumptions of the ideal gas model.
- 9. Calculate pressure of 0.20 mol gas in 4.0 dm³ at 300 K using R = 8.314 kPa $dm^3\ mol^{-1}\ K^{-1}$.
- 10. Define activation energy and explain effective collisions.
- 11. Explain effects of concentration, temperature and surface area on reaction rate.
- 12. Explain why a catalyst speeds a reaction but is not consumed overall.
- 13. A reaction loses 2.5 g in 50 s. Calculate average mass-loss rate.
- 14. Plan a fair investigation of the effect of temperature on reaction rate.
- 15. Explain two practical reasons for controlling reaction rates.

Glossary

Aqueous (aq): Dissolved in water.

Anion: A negatively charged ion.

Cation: A positively charged ion.

Catalyst: A substance that increases reaction rate by providing an alternative pathway and is regenerated overall.

Concentration: Amount of solute per unit volume of solution.

Crystallisation: Formation of crystals from solution.

Effective collision: A collision with suitable orientation and enough energy to form products.

Activation energy: Minimum energy required for a reaction pathway to proceed.

Molar gas volume: Volume occupied by one mole of gas at stated temperature and pressure.

Net ionic equation: Equation showing only species undergoing chemical change.

Precipitate: An insoluble solid formed from solution.

Rate of reaction: Change in measured reactant or product quantity per unit time.

Solubility: Maximum amount of solute that dissolves under stated conditions.

Spectator ion: Ion unchanged during a reaction and omitted from the net ionic equation.

Water of crystallisation: Water incorporated in a definite ratio into a crystalline salt.

Final Summary

Physical Chemistry links observable changes to particle behaviour and measurable quantities. Salt chemistry develops skills in naming compounds, balancing formulae, predicting solubility, preparing pure samples and writing ionic equations. Gas laws explain how pressure, volume, temperature and amount of gas are related; successful calculations require the correct relationship, kelvin temperatures and consistent units. Reaction-rate chemistry explains why conditions alter the frequency and effectiveness of collisions, how catalysts lower activation energy, and how graphs and fair investigations compare rates. These ideas apply in laboratory practice, manufacturing, environmental management and everyday technology.

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How to Use These Notes

These notes introduce the core ideas of hydrocarbons through explanations, structural formulae, worked examples, revision questions and practical safety guidance. Read the explanation, study each formula, then attempt the questions without looking at the answers. Use the yellow-highlighted illustration prompts as instructions for inserting clear, labelled textbook-style diagrams.

Learning Outcomes

- Classify hydrocarbons as alkanes, alkenes or alkynes using bonding and formulae.
- Represent hydrocarbons with molecular, condensed, structural and displayed formulae.
- Apply basic systematic naming rules to straight-chain and simple branched hydrocarbons.
- Explain structural isomerism and simple geometrical isomerism in alkenes.
- Write and balance equations for combustion and selected addition reactions.
- Explain substitution, addition and addition polymerisation.
- Describe uses of hydrocarbons and consider safety and environmental issues.

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1. Introduction to Organic Chemistry

Organic chemistry is the study of carbon compounds, especially compounds containing carbon–carbon and carbon–hydrogen bonds. Carbon forms a very large number of compounds because each carbon atom can form four covalent bonds and carbon atoms can bond to one another to form straight chains, branched chains and rings. These structures may contain single, double or triple bonds.

Hydrocarbons are the simplest organic compounds because they contain only carbon and hydrogen. They are important as fuels, raw materials for plastics and synthetic fibres, lubricants, solvents and many other products. Natural gas, petrol, kerosene, diesel and paraffin wax contain mixtures of hydrocarbons. The composition of a fuel affects its volatility, ignition behaviour, energy release and environmental impact.

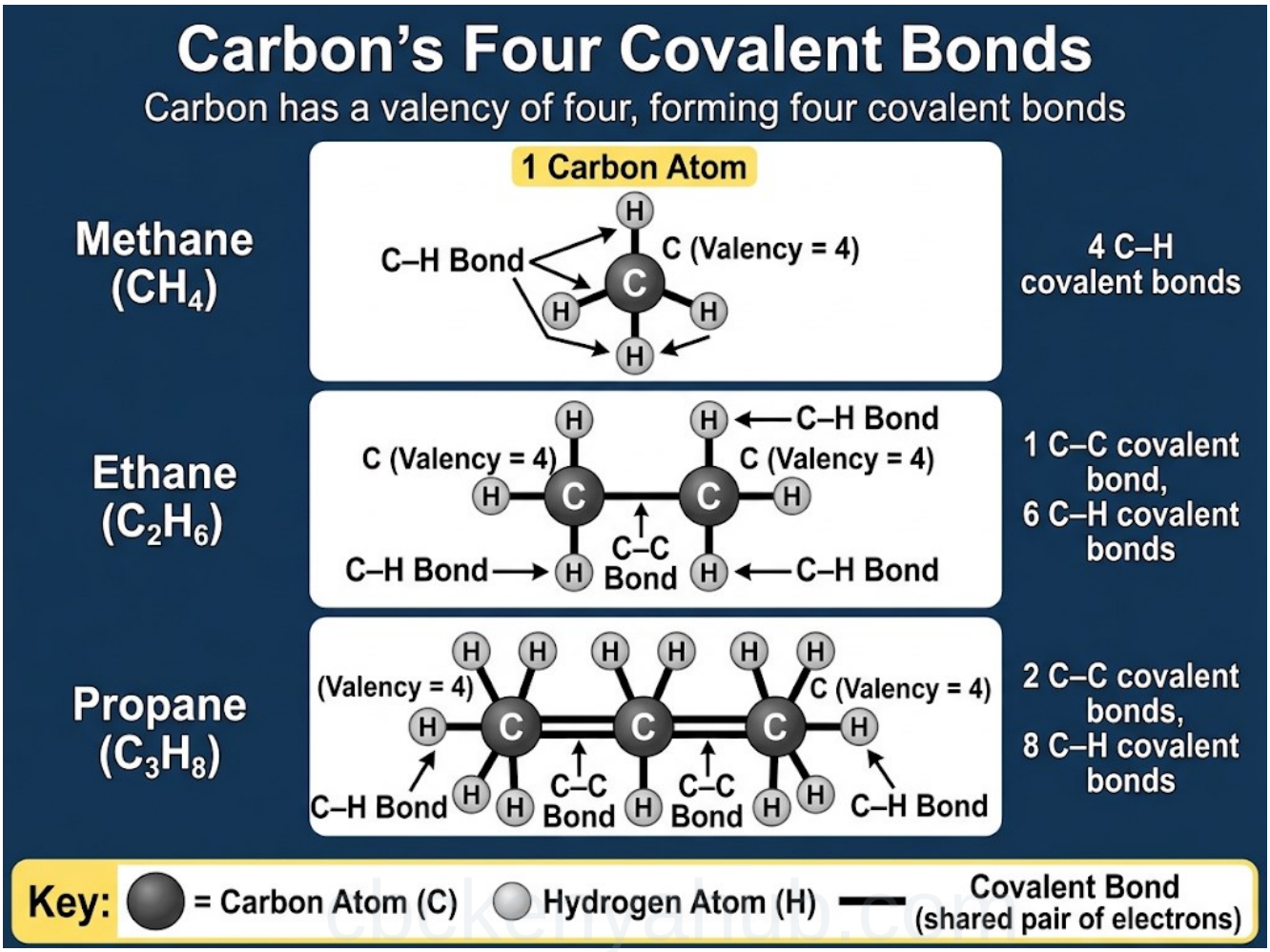
Organic compounds are commonly represented using molecular formulae, structural formulae, displayed formulae, condensed formulae and skeletal formulae. A molecular formula gives the actual number of each type of atom in a molecule. A structural formula shows how atoms are connected. A displayed formula shows every covalent bond. Different representations are useful for different tasks, but they must all describe the same molecule.

Carbon atoms are tetravalent: a carbon atom normally forms four covalent bonds. Hydrogen forms one covalent bond. In a hydrocarbon, the total number of bonds around each carbon must be four. This rule helps learners draw structures and identify errors. A carbon atom with five ordinary covalent bonds in a simple displayed formula signals an incorrect structure.

Organic chemistry also requires careful laboratory practice. Many hydrocarbons are volatile and flammable; vapours may ignite at a distance from the liquid. Use small quantities, keep containers closed, avoid naked flames unless the method specifically requires a controlled flame, and work with suitable ventilation. Never identify an unknown hydrocarbon by directly inhaling it.

Check Your Understanding

- Define organic chemistry and hydrocarbons.
- Explain why carbon forms many compounds.
- Distinguish molecular, structural and displayed formulae.
- State two safety precautions when handling hydrocarbons.



2. Hydrocarbons: Classification

Hydrocarbons can be classified according to the types of carbon–carbon bonds present and the arrangement of their carbon skeletons. Acyclic hydrocarbons have open chains; these chains may be straight or branched. Cyclic hydrocarbons contain rings. At this level, the most important open-chain families are alkanes, alkenes and alkynes.

Alkanes are saturated hydrocarbons: they contain carbon–carbon single bonds only. Their open-chain general formula is C_nH_{2n+2}. Examples include methane, ethane, propane and butane. Saturated means that the carbon skeleton has the maximum number of hydrogen atoms possible for an open chain with single bonds.

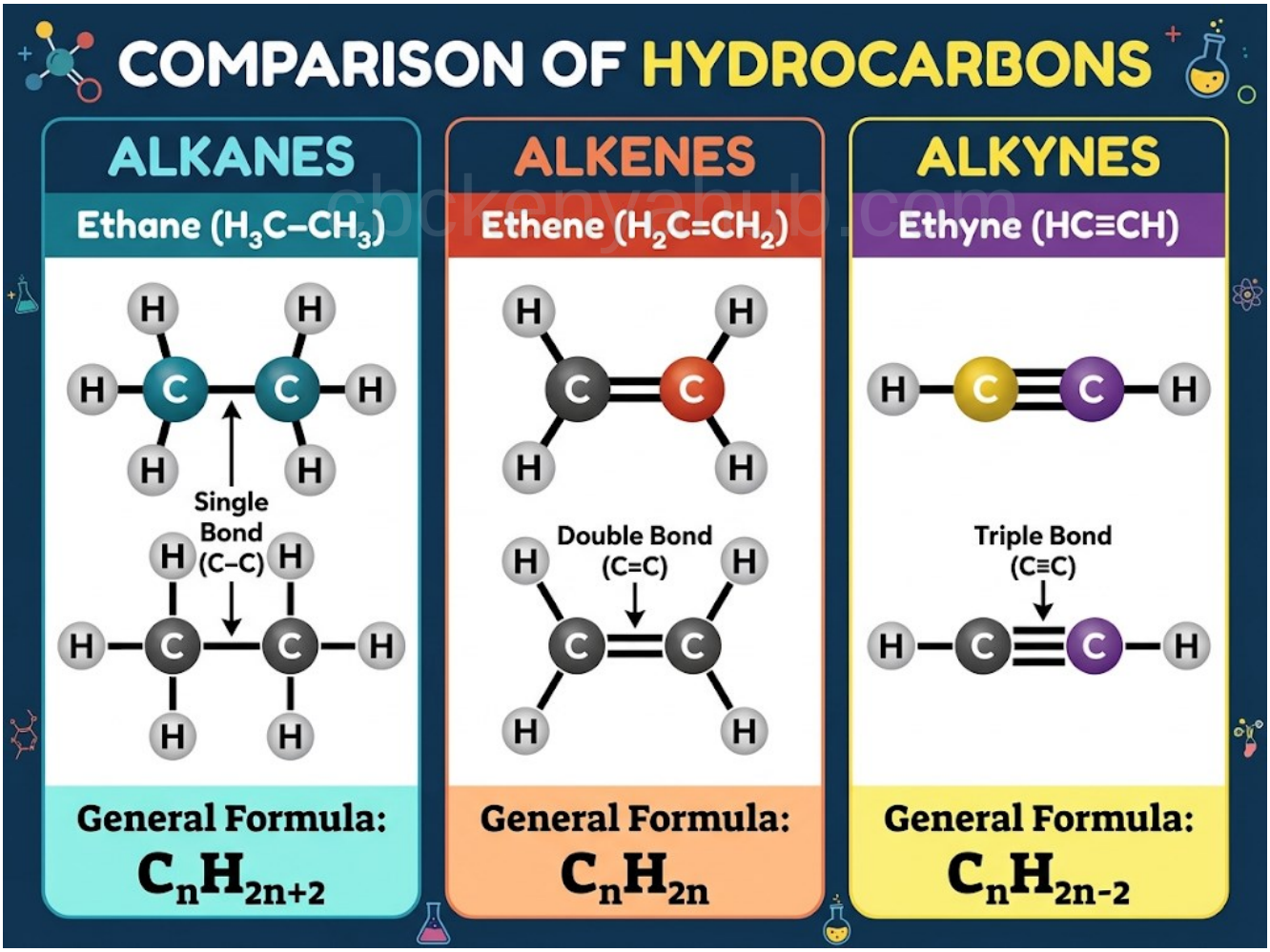
Alkenes are unsaturated hydrocarbons containing at least one carbon–carbon double bond. A simple open-chain alkene with one double bond has the general formula C_nH_{2n}. Ethene (C_2H_4) and propene (C_3H_6) are examples. Alkynes contain at least one carbon–carbon triple bond; a simple open-chain alkyne with one triple bond has the formula C_nH_{2n-2}. Ethyne (C_2H_2) is the simplest alkyne.

These general formulae apply only when the stated assumptions are satisfied: open-chain compounds, one double bond for the alkene formula, or one triple bond for the alkyne formula. Rings and additional multiple bonds alter the hydrogen count. Always inspect the structure before applying a formula.

Unsaturation usually makes alkenes and alkynes more reactive in addition reactions than alkanes. The multiple bond can be converted into single bonds as atoms or groups add across it. This difference provides a practical way to distinguish families in appropriate tests.

Check Your Understanding

- Classify a hydrocarbon as an alkane, alkene or alkyne from its structure.
- Write the general formulae for the three open-chain families.
- Explain the terms saturated and unsaturated.
- State why general formulae must be applied carefully.



3. Carbon Bonding and Structural Formulae

A covalent bond forms when atoms share electron pairs. A single carbon–carbon bond contains one shared pair; a double bond contains two shared pairs; and a triple bond contains three shared pairs. In simple structural drawings, single bonds are shown by one line, double bonds by two parallel lines and triple bonds by three parallel lines.

Condensed structural formulae compress a displayed structure into groups. Ethane may be written CH_3CH_3, propane as $CH_3CH_2CH_3$ and butane as $CH_3CH_2CH_2CH_3$. For ethene, $CH_2{=}CH_2$ makes the double bond explicit. For ethyne, HC≡CH shows the triple bond. The condensed formula must still communicate the position of multiple bonds unambiguously.

To check a displayed structure, count the bonds around every carbon atom. For example, in ethene each carbon forms a double bond to the other carbon and two single bonds to hydrogen, giving four bonds around each carbon. In ethyne, each carbon forms a triple bond to the other carbon and one single bond to hydrogen.

Structural formulae matter because compounds with the same molecular formula can have different atom connections. Such compounds are structural isomers. For example, butane and 2-methylpropane both have molecular formula C_4H_{10}, but one has a continuous four-carbon chain while the other has a branched three-carbon main chain with a methyl branch.

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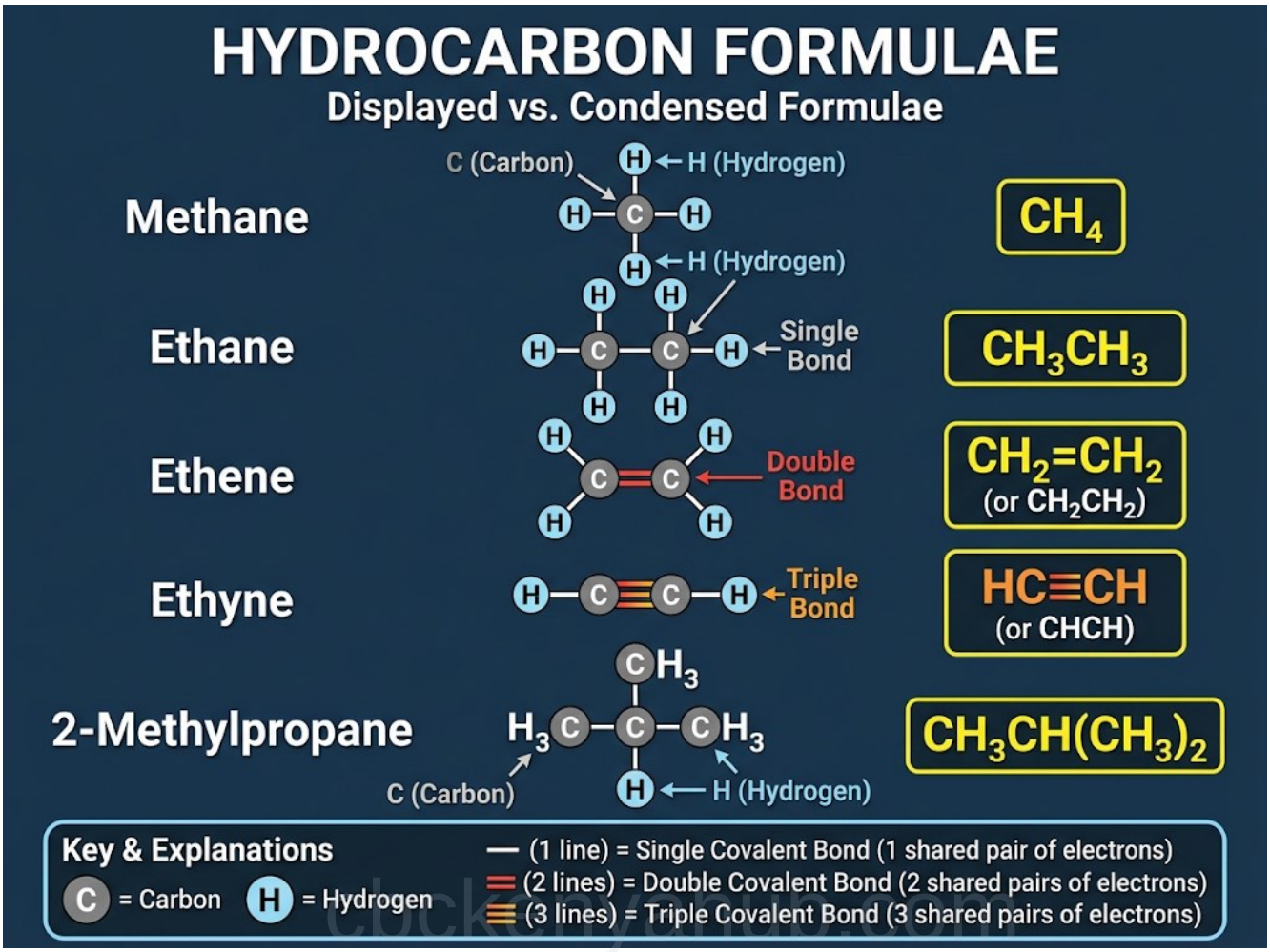
Models can help learners visualise structures, but models are representations rather than exact pictures of molecules. Ball-and-stick models emphasise bonds and angles; space-filling models give a better impression of relative atomic sizes and the space occupied by a molecule.

Check Your Understanding

- Write condensed formulae for ethane, propane and butane.
- Explain how a displayed formula can be checked for carbon valency.
- Define structural isomers.
- Give one strength and one limitation of molecular models.



4. Naming Hydrocarbons: Carbon-Chain Roots

Systematic names help chemists communicate the exact structure of a compound. For simple hydrocarbons, the root indicates the number of carbon atoms in the parent chain. The first ten roots are meth- (1), eth- (2), prop- (3), but- (4), pent- (5), hex- (6), hept- (7), oct- (8), non- (9) and dec- (10).

The ending indicates the family: -ane for alkanes, -ene for alkenes and -yne for alkynes. Thus, a two-carbon saturated hydrocarbon is ethane; a two-carbon hydrocarbon with a double bond is ethene; and a two-carbon hydrocarbon with a triple bond is ethyne. The number of carbon atoms must be counted from the chosen parent chain, not from the total number of carbon atoms in every branch.

Naming begins by identifying the longest continuous carbon chain that contains the required multiple bond, if one is present. For alkenes and alkynes, the parent chain must include the double or triple bond. A chain that is longer but excludes the multiple bond may not be the correct parent chain.

Branches are named as substituents. A one-carbon branch is methyl (–CH_3), a two-carbon branch is ethyl (–CH_2CH_3), and a three-carbon branch is propyl. The name includes a number showing the carbon atom to which each branch is attached. Use the lowest appropriate number set while following the rules for the family.

Common names remain widely used for some substances, such as acetylene for ethyne and ethylene for ethene. In school naming exercises, use systematic names unless the question requests a common name.

Check Your Understanding

- Recall the roots for one to ten carbon atoms.
- Match -ane, -ene and -yne to the correct family.
- Explain how the parent chain is selected.
- Name the substituent formed from one carbon atom.

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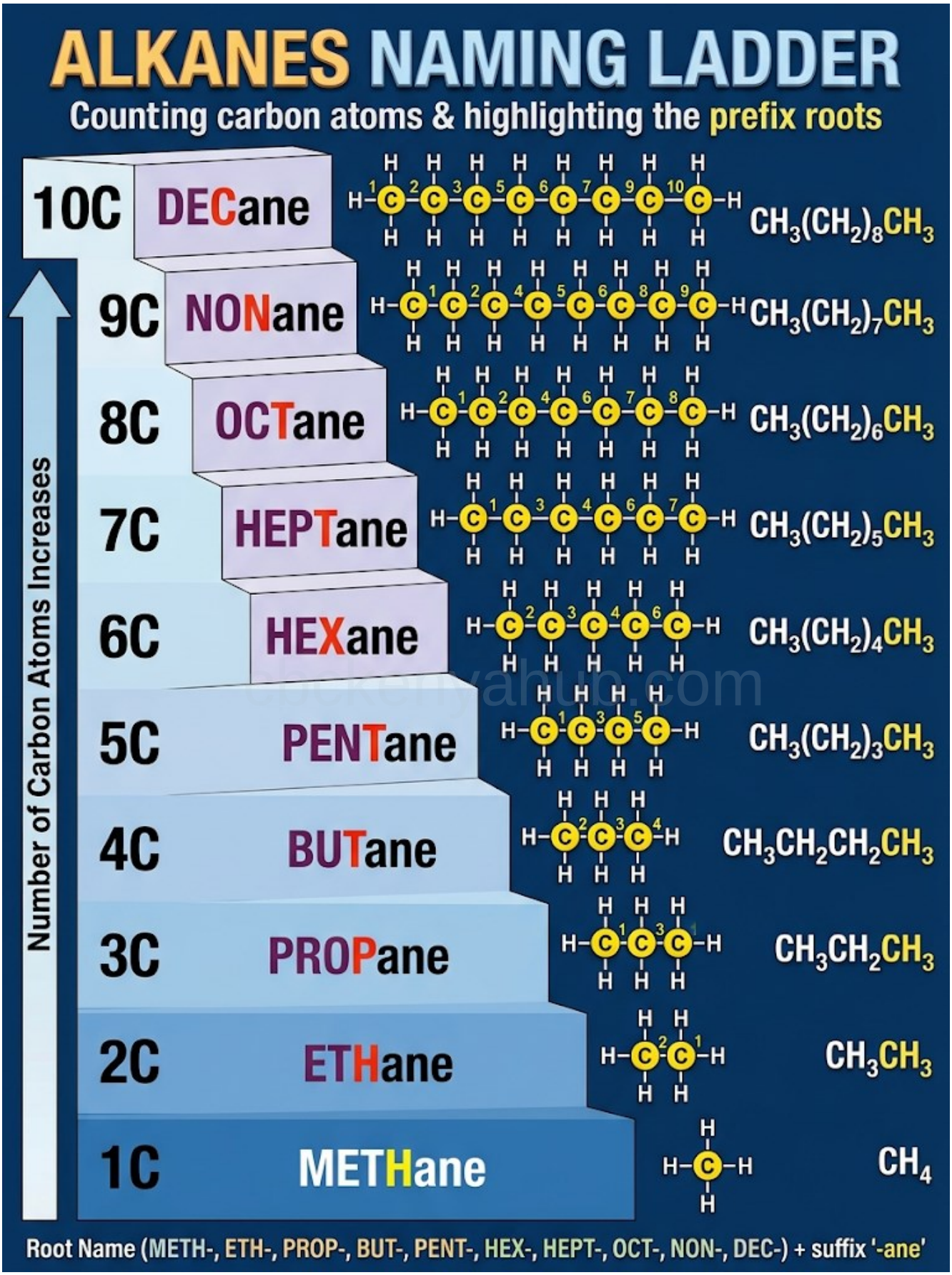

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5. Naming Alkanes

To name a simple alkane, identify the longest continuous carbon chain and select the correct root. Since the compound contains only single bonds, use the ending -ane. If the chain is straight and unbranched, its name is straightforward: $CH_3CH_2CH_2CH_3$ is butane because it contains four carbon atoms.

For a branched alkane, select the longest continuous chain as the parent. Number the chain from the end that gives the branch or branches the lowest possible number. Identify each branch and its position. A methyl branch on carbon 2 of a five-carbon chain gives 2-methylpentane. If identical branches occur more than once, use di-, tri- or tetra- and list every position.

When different substituents are present, list them alphabetically in the name; prefixes such as di- and tri- are generally not used when alphabetising substituent names. Separate numbers with commas and separate numbers from words with hyphens. Do not insert spaces in the name.

Worked example: Name $CH_3–CH(CH_3)–CH_2–CH_3$. The longest continuous chain contains four carbon atoms, so the parent is butane. Numbering from the nearer end places the methyl branch at carbon 2 rather than carbon 3. The name is 2-methylbutane.

Common errors include choosing a chain that is not the longest, numbering from the wrong end, forgetting a branch locant and using a number without a hyphen. Check the parent chain first, then numbering, then substituents, then punctuation.

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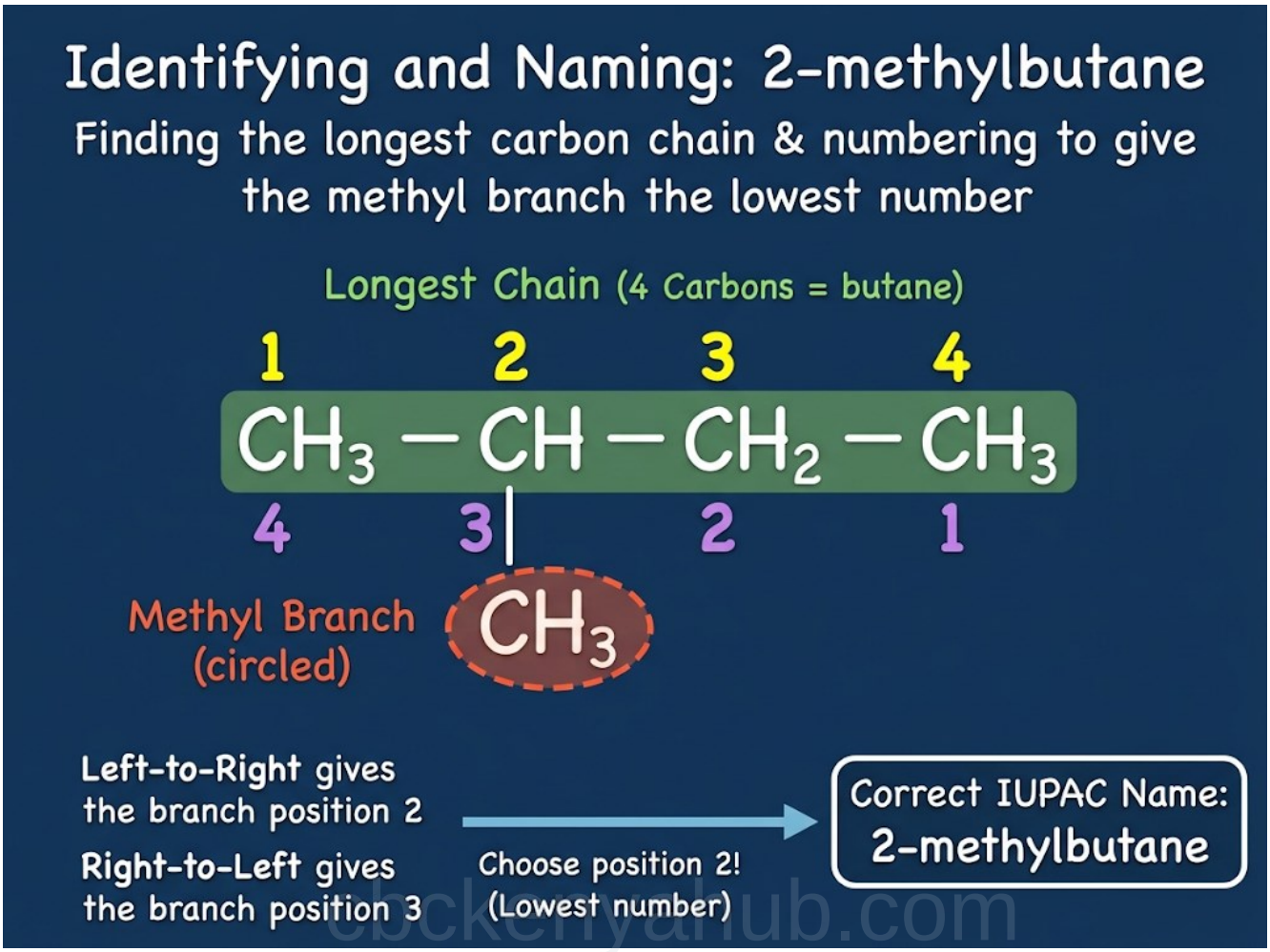
Check Your Understanding

- Name 2-methylpropane from its structure.
- Explain why a four-carbon parent chain takes precedence over a shorter chain.
- Name $CH_3CH(CH_3)CH_2CH_2CH_3$.
- Identify two common naming errors.



6. Isomerism in Alkanes

Structural isomers have the same molecular formula but different arrangements of atoms. As the number of carbon atoms increases, the number of possible carbon skeletons generally increases. This is one reason organic chemistry contains many distinct compounds.

Butane, C_4H_{10}, has two structural isomers: butane and 2-methylpropane. Both contain four carbon atoms and ten hydrogen atoms. In butane the carbon atoms form an unbranched chain; in 2-methylpropane the carbon skeleton is branched. Their physical properties differ because molecular shape and surface contact influence intermolecular attractions.

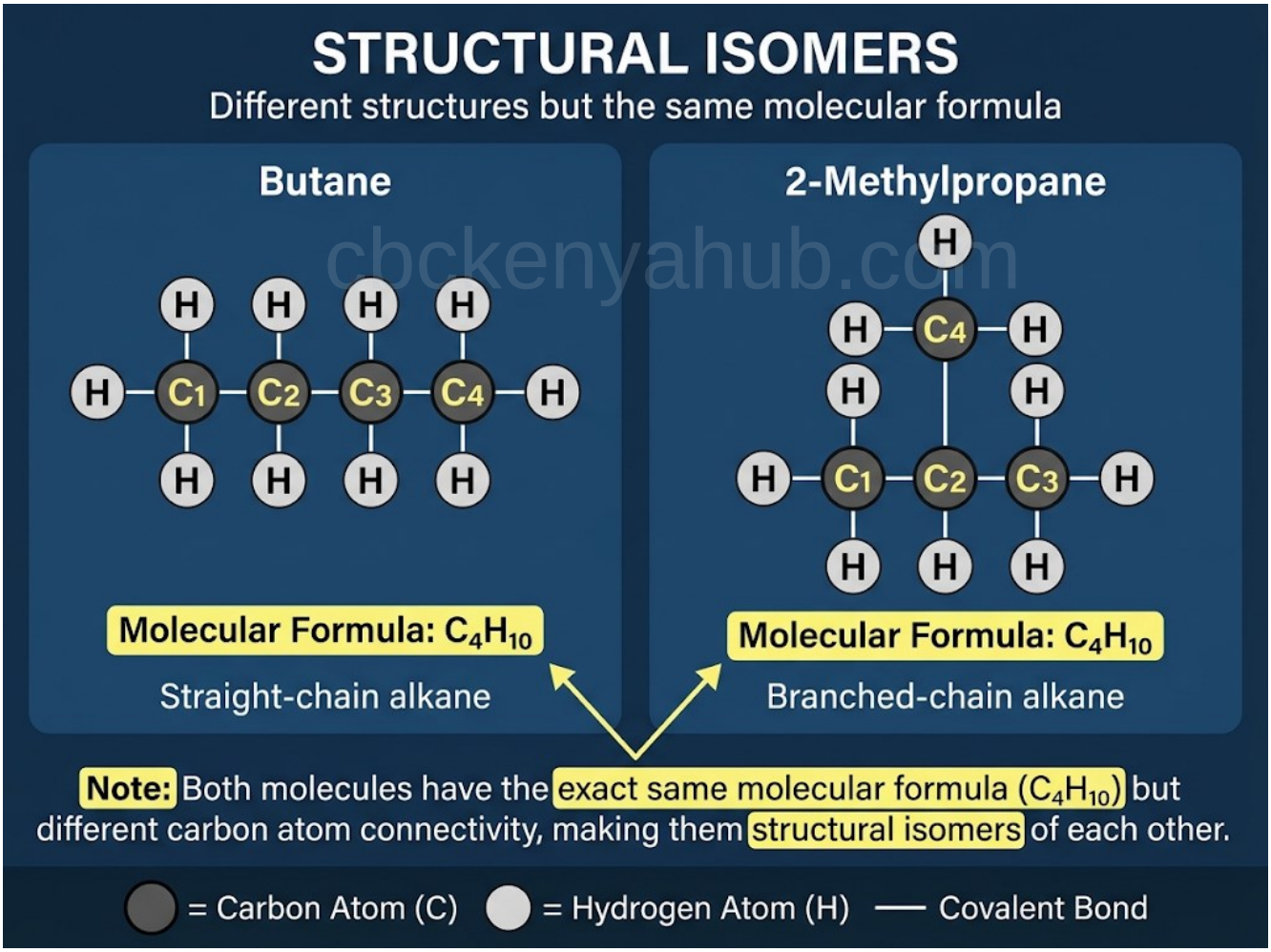
Pentane, C_5H_{12}, has three structural isomers: pentane, 2-methylbutane and 2,2-dimethylpropane. A reliable method for drawing isomers is to start with the longest possible chain, then shorten the parent chain and place branches in unique positions. Avoid drawing the same structure in a different orientation and counting it as a new isomer.

Isomer counting should preserve the molecular formula and valid carbon valency. For example, a carbon atom cannot be assigned five ordinary covalent bonds. Check that every proposed structure has the correct number of carbon atoms and hydrogen atoms.

Isomerism has practical significance. Different isomers may have different boiling points, volatility, ignition behaviour and uses. In fuel chemistry, branching often changes combustion characteristics, which is one reason the structure of hydrocarbon mixtures matters.

Check Your Understanding

- Define structural isomerism.
- Draw the two isomers of C_4H_{10}.
- List the three structural isomers of C_5H_{12}.
- Describe a systematic approach to finding isomers.



7. Physical Properties of Alkanes

Alkanes are generally non-polar molecules. Their molecules attract one another mainly through London dispersion forces. These attractions are relatively weak compared with covalent bonds within the molecules, so boiling an alkane separates molecules without breaking the C–C and C–H covalent bonds.

For a homologous series of straight-chain alkanes, boiling point generally increases as the number of carbon atoms increases. Larger molecules have more electrons and greater surface contact, leading to stronger dispersion attractions. More energy is therefore needed to separate the molecules.

Small alkanes such as methane, ethane, propane and butane are gases at room temperature under ordinary pressure. Medium-sized alkanes are commonly liquids, while much larger alkanes can be waxy solids. These descriptions are approximate because physical state depends on temperature and pressure.

Alkanes are insoluble or only very slightly soluble in water because they are non-polar, while water is polar. They are more soluble in many non-polar organic solvents. Most alkanes are less dense than water, so a separate alkane layer generally floats on water, although safe laboratory practice should always guide handling.

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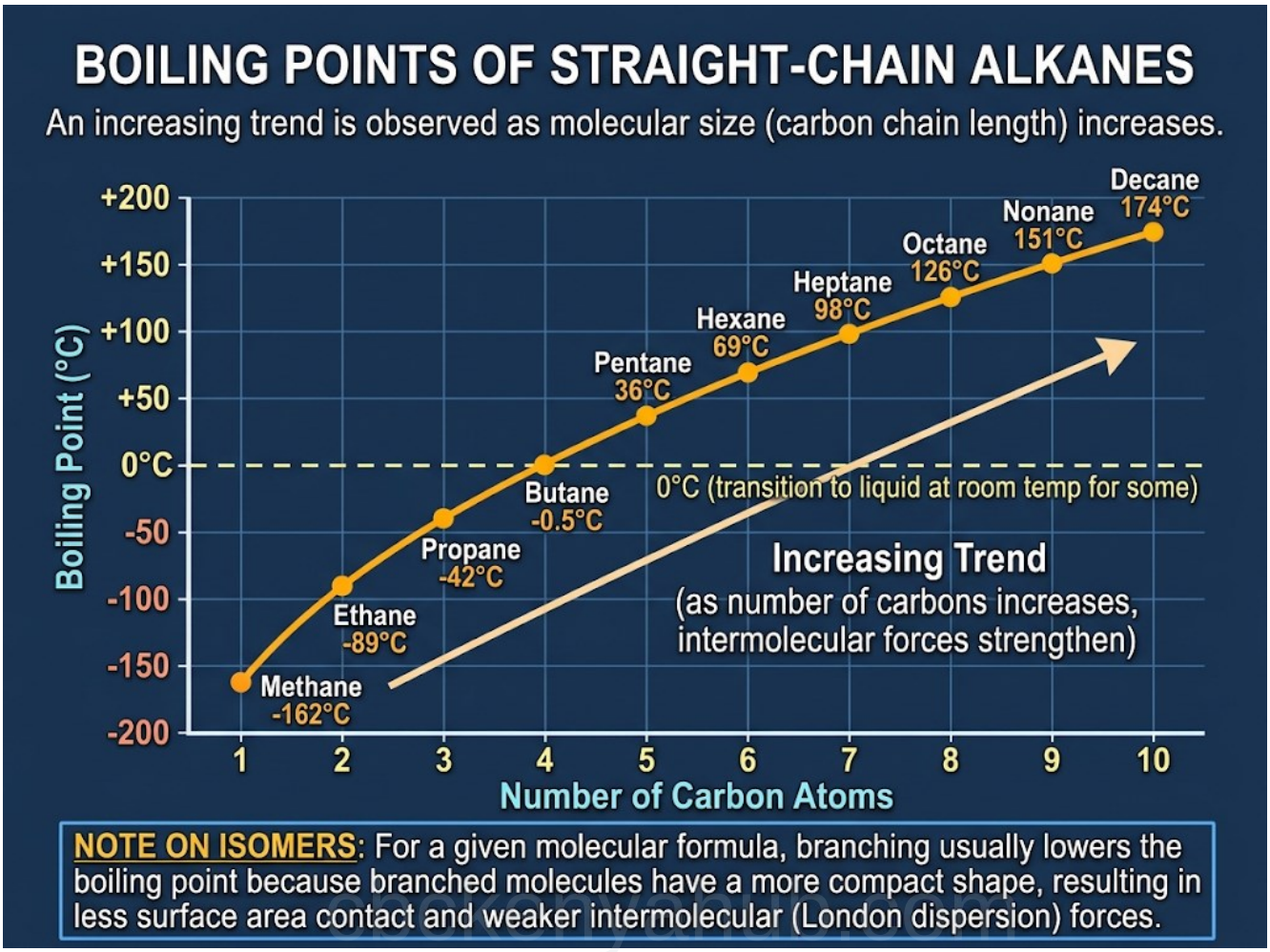
Branching affects boiling point. For isomers with the same molecular formula, a more compact, highly branched molecule often has a lower boiling point than a straight-chain isomer because it has less surface contact. This is a useful trend, not a substitute for checking measured data.

Check Your Understanding

- Explain the trend in boiling points down the alkane series.
- Why are alkanes poorly soluble in water?
- State the typical physical states of short- and long-chain alkanes.
- Explain how branching can affect boiling point.



8. Chemical Properties of Alkanes: Combustion

Combustion is a reaction with oxygen that releases energy. Complete combustion occurs when a hydrocarbon burns in sufficient oxygen, producing carbon dioxide and water. For example: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. The equation is balanced because there is one carbon atom, four hydrogen atoms and four oxygen atoms on each side.

To balance a hydrocarbon combustion equation, first balance carbon by placing the same number of CO_2 molecules as carbon atoms in the fuel. Next balance hydrogen using H_2O. Finally balance oxygen using O_2. If an odd number of oxygen atoms is required, a fractional coefficient may be used temporarily, then all coefficients can be multiplied to obtain whole numbers.

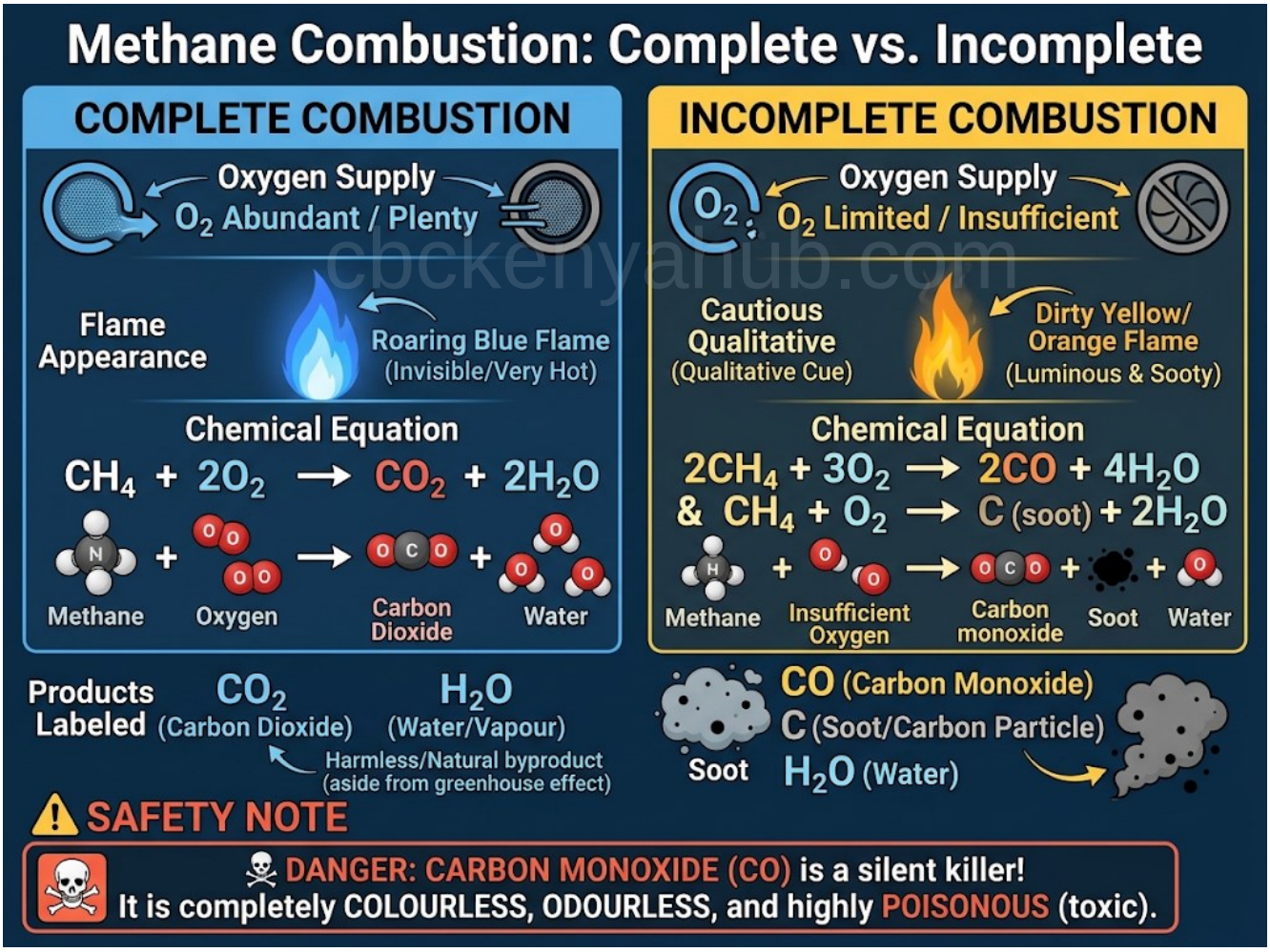
Worked example: Balance $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$. Three carbon atoms require $3CO_2$. Eight hydrogen atoms require $4H_2O$. The products contain $6 + 4 = 10$ oxygen atoms, so $5O_2$ is needed. The balanced equation is $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$.

In limited oxygen, incomplete combustion may produce carbon monoxide and/or carbon (soot), as well as water. Carbon monoxide is highly poisonous because it interferes with oxygen transport in the blood. Soot contributes to air pollution and can affect respiratory health.

Combustion is useful in heating, transport and electricity generation, but it also releases carbon dioxide, a greenhouse gas. Safe appliances require adequate ventilation, correct maintenance and suitable exhaust systems. Never operate a fuel-burning appliance in a poorly ventilated space.

Check Your Understanding

- Balance the complete combustion of ethane.
- Distinguish complete from incomplete combustion.
- Explain why carbon monoxide is dangerous.
- State one benefit and two environmental concerns associated with hydrocarbon combustion.



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9. Substitution Reactions of Alkanes

Alkanes are relatively unreactive under ordinary conditions because their C–C and C–H bonds are strong and non-polar. One characteristic reaction is substitution with a halogen, such as chlorine, in the presence of ultraviolet light. In substitution, one atom in the original molecule is replaced by another atom.

For example, methane reacts with chlorine under suitable photochemical conditions: $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$. The products are chloromethane and hydrogen chloride. The reaction may continue, producing more highly substituted products if chlorine remains available and conditions permit.

Ultraviolet light supplies energy to initiate the reaction by splitting some chlorine molecules into reactive chlorine radicals. At school level, the key observable idea is that light is needed to initiate the reaction; detailed radical mechanisms may be treated as an extension where appropriate.

Substitution differs from addition. In substitution, one atom or group replaces another. In addition, atoms or groups add across a multiple bond, and the multiple bond becomes a single bond. This distinction is important when comparing alkane and alkene chemistry.

Halogenated hydrocarbons have industrial uses but some are toxic or environmentally harmful. Do not attempt halogenation demonstrations without approved procedures, proper ventilation, risk assessment and teacher supervision. The equations in these notes are for learning, not instructions to carry out an unsupervised reaction.

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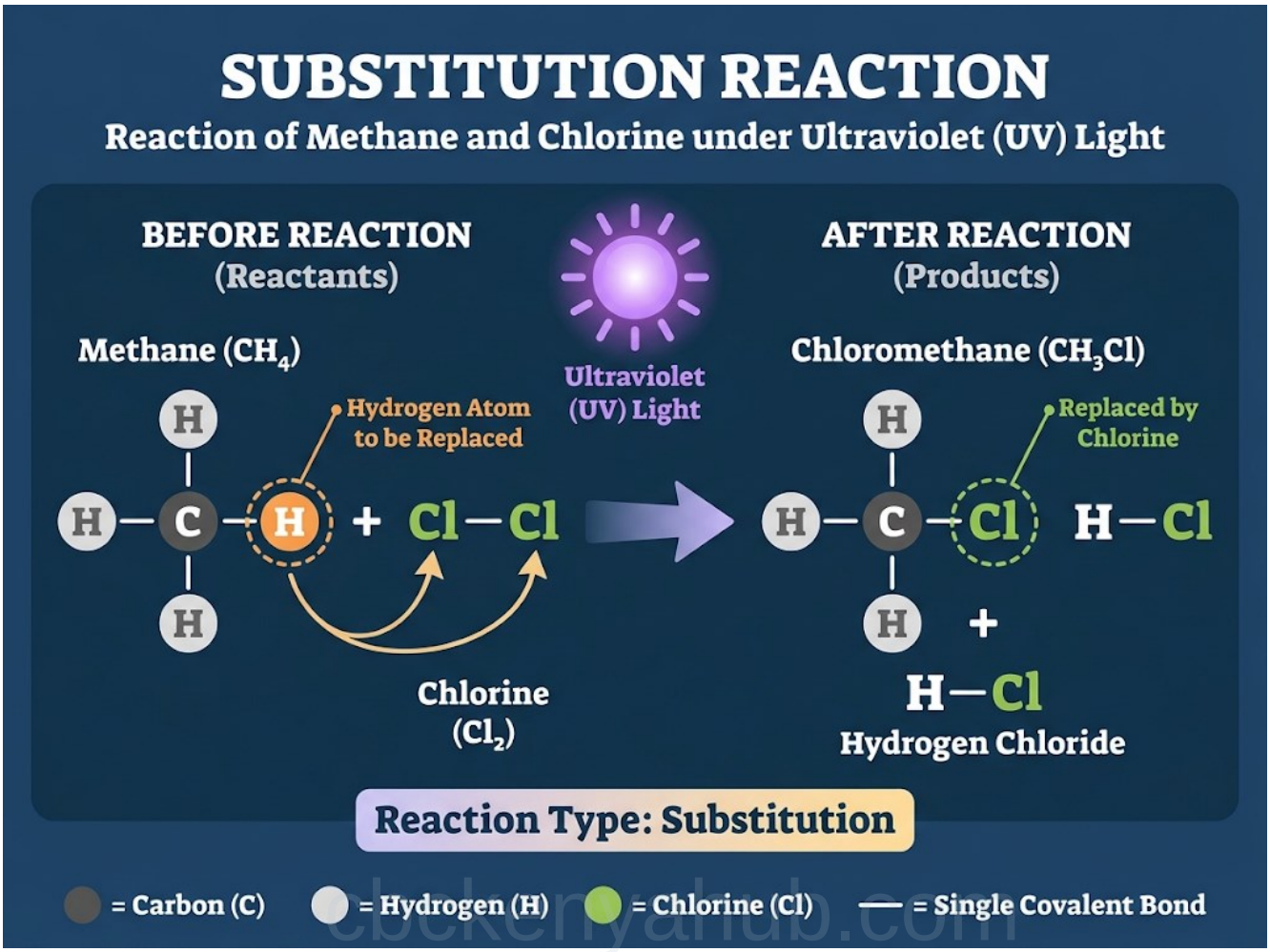
Check Your Understanding

- Define substitution reaction.
- Write the equation for methane reacting with chlorine.
- State the role of ultraviolet light.
- Distinguish substitution from addition.



10. Alkenes: Structure and General Formula

Alkenes are unsaturated hydrocarbons containing a carbon–carbon double bond. The simplest alkene is ethene, C_2H_4, with structural formula $CH_2{=}CH_2$. The double bond consists of two shared electron pairs and prevents free rotation in the same way as a single bond permits rotation.

The general formula for a simple open-chain alkene with one carbon–carbon double bond is C_nH_{2n}. For example, propene is C_3H_6 and butene is C_4H_8. This formula should not be applied blindly to cyclic compounds or molecules containing additional multiple bonds.

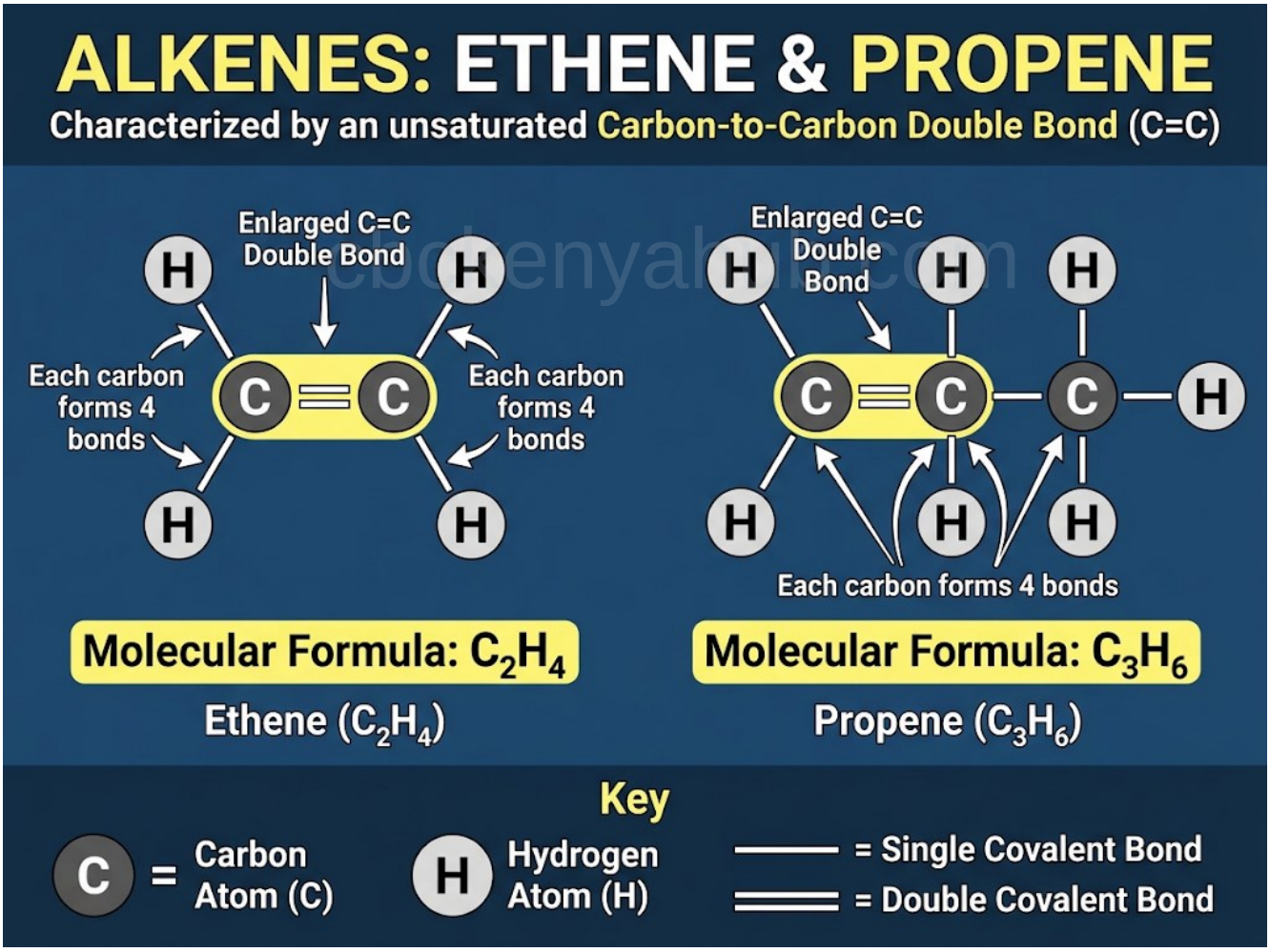
An alkene must contain at least two carbon atoms because a carbon–carbon double bond requires two carbon atoms. Alkenes with four or more carbon atoms may have the double bond in different positions or may have branched skeletons, creating structural isomers.

Alkenes are important industrial raw materials. Ethene is used to manufacture poly(ethene) and is involved in fruit ripening. Propene is used to make poly(propene). The double bond allows alkenes to undergo addition reactions that convert small molecules into products with new bonds.

To recognise an alkene in a structural formula, look for C=C. A formula containing only C–C single bonds belongs to a different structure unless a ring or another feature changes the classification. Always check the displayed or condensed formula rather than relying on the name.

Check Your Understanding

- Define an alkene.
- Write the general formula for a simple open-chain alkene with one double bond.
- Explain why ethene has two carbon atoms.
- Give two uses of alkenes.



11. Naming Alkenes

To name an alkene, choose the longest continuous carbon chain that contains the double bond. Use the appropriate carbon-chain root and the ending -ene. Number the parent chain from the end nearest the double bond so that the double bond receives the lowest possible locant.

For a double bond beginning at carbon 1 in a four-carbon chain, the name is but-1-ene. If the double bond begins at carbon 2, the name is but-2-ene. The locant indicates the first carbon atom in the double bond. Use consistent school-level IUPAC style, for example but-1-ene and pent-2-ene.

Branches are included as for alkanes. The parent chain must include the double bond, and numbering prioritises giving the double bond the lowest possible number before branch positions are considered. For example, a methyl branch on carbon 3 of but-1-ene is named 3-methylbut-1-ene when that structure is valid.

Worked example: Name $CH_3–CH=CH–CH_3$. The longest chain has four carbon atoms, so the root is but-. The double bond is between carbon 2 and carbon 3 whichever end is used. The name is but-2-ene.

Check the position of the double bond, carbon count, branch positions and punctuation. A common error is to choose a longer chain that does not contain the double bond. Another is to number from the end farthest from the double bond.

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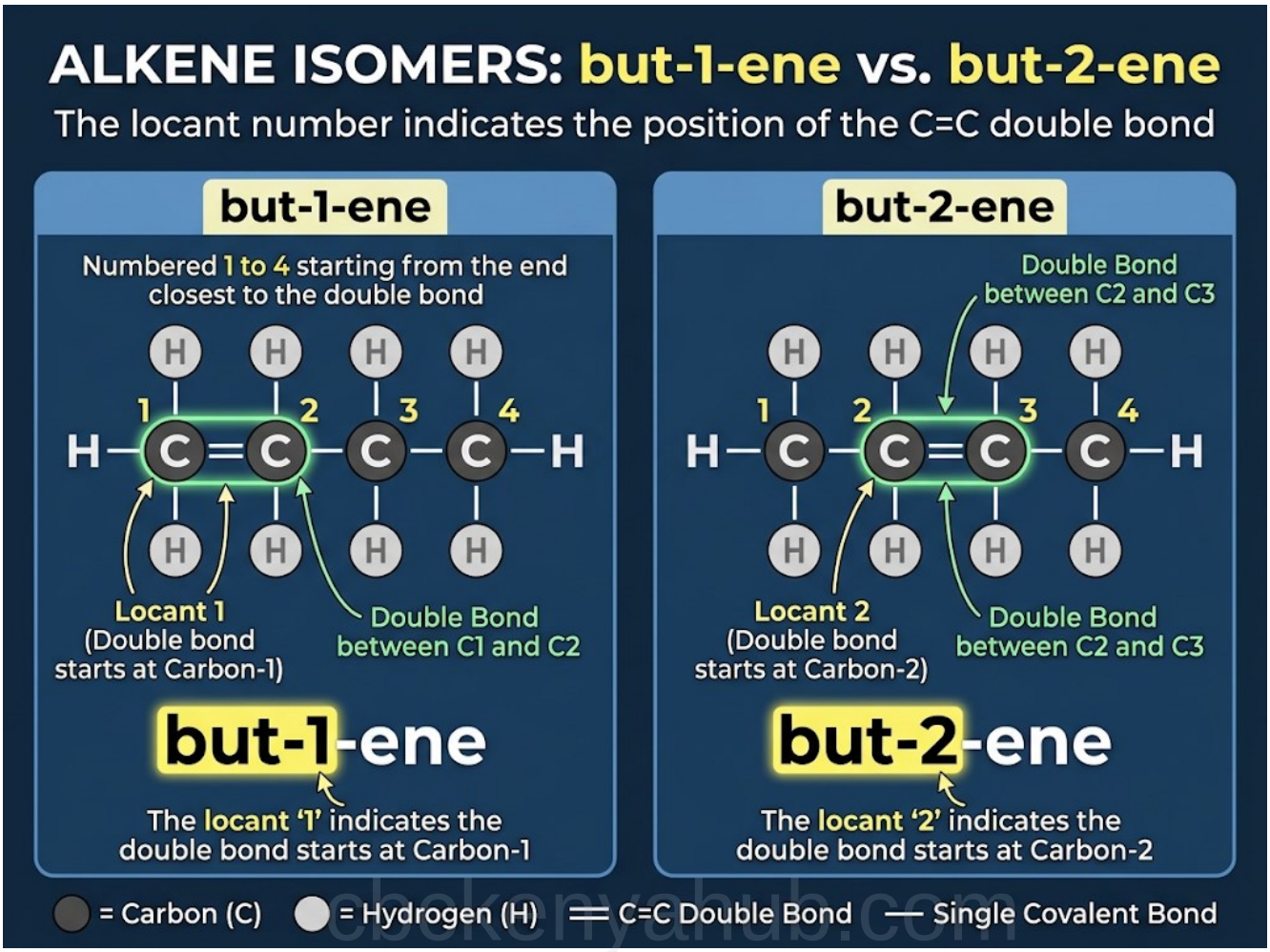
Check Your Understanding

- Name $CH_2=CH–CH_2–CH_3$.
- Name $CH_3–CH=CH–CH_3$.
- Explain why the parent chain must contain the double bond.
- State how the double-bond position is indicated.



12. Isomerism in Alkenes

Alkenes can exhibit structural isomerism because the carbon skeleton may be branched and because the double bond can occupy different positions. For example, but-1-ene and but-2-ene have molecular formula C_4H_8 but different double-bond positions. 2-methylpropene is another structural isomer of C_4H_8.

Geometrical isomerism can occur around a carbon–carbon double bond because rotation is restricted. For a pair of geometrical isomers to exist, each carbon of the double bond must be attached to two different groups. In but-2-ene, each double-bond carbon is attached to hydrogen and a methyl group, so two arrangements are possible.

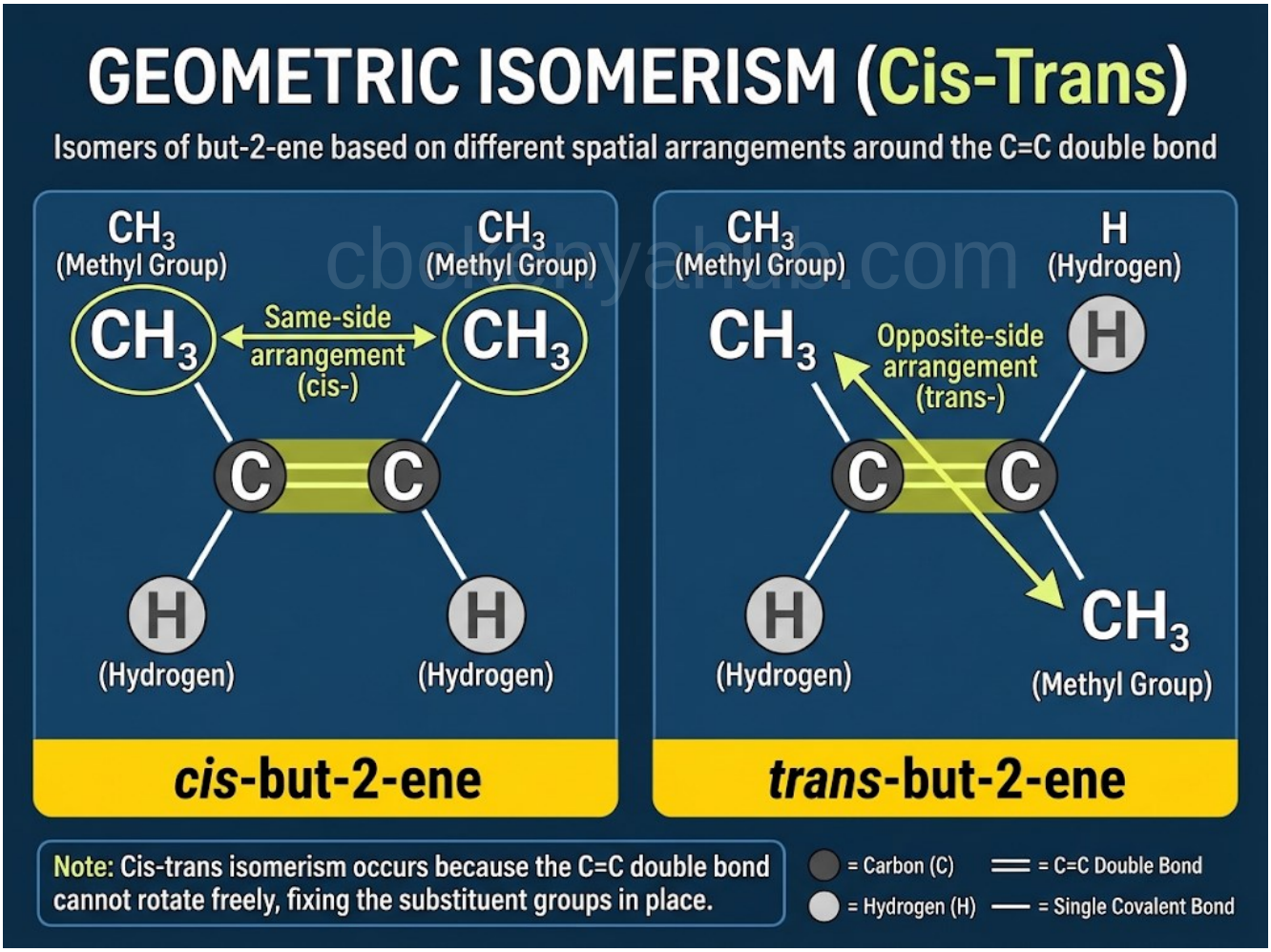
In cis-but-2-ene, the two methyl groups lie on the same side of the double bond; in trans-but-2-ene, they lie on opposite sides. These arrangements have the same connectivity but different spatial arrangement. They can have different physical properties.

Do not assume every alkene has cis–trans isomers. In but-1-ene, one double-bond carbon has two hydrogen atoms, so the required pair of different groups is absent on that carbon. Therefore, the usual cis–trans distinction does not apply.

When drawing isomers, keep carbon valency correct and ensure the formula remains unchanged. For geometrical isomers, show the double bond clearly and place the groups on the correct sides. A simple two-dimensional drawing can communicate the distinction if it is carefully labelled.

Check Your Understanding

- Give three structural isomers of C_4H_8 that are alkenes.
- Explain why rotation around a double bond is restricted.
- Distinguish cis-but-2-ene from trans-but-2-ene.
- Why does but-1-ene not show the usual cis–trans isomerism?



13. Chemical Test for Unsaturation

A common qualitative test for carbon–carbon double bonds uses bromine water. Under suitable classroom conditions, an alkene decolourises orange-brown bromine water as bromine adds across the double bond. The reaction converts an unsaturated compound into a saturated dibromoalkane.

For ethene, a simplified equation is $CH_2=CH_2 + Br_2 \rightarrow CH_2Br–CH_2Br$. The bromine atoms add to the two carbon atoms that were joined by the double bond. This is an addition reaction. The observation is a colour change from orange-brown to colourless when enough alkene reacts.

Alkanes generally do not decolourise bromine water rapidly in the absence of ultraviolet light or other suitable conditions. This difference can help distinguish an alkane from an alkene, but a test must be interpreted cautiously: other reducing substances or reactive compounds may also affect bromine water.

Use small quantities, wear eye protection, avoid skin contact and follow the teacher's instructions for chemical disposal. Bromine solutions are hazardous and should not be handled casually. The test should be conducted only in an approved educational setting with the appropriate safety controls.

A good practical record includes the aim, apparatus, method, observations, conclusion and safety precautions. Record what is actually observed, not what you expected to see. If the colour does not disappear, discuss possible explanations such as insufficient sample, unsuitable conditions or an incorrect sample.

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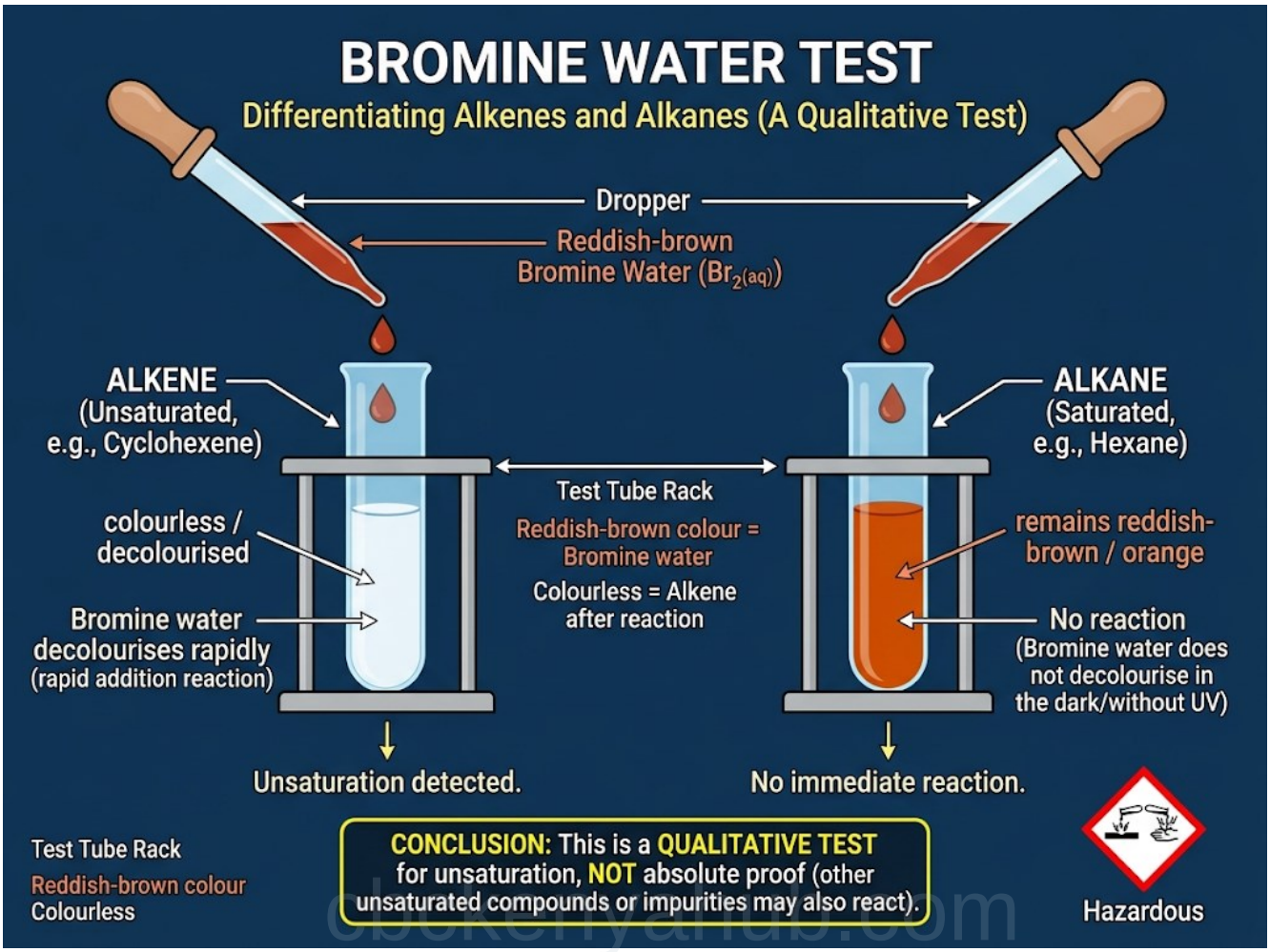
Check Your Understanding

- Describe the bromine-water test for an alkene.
- Write the addition equation for ethene and bromine.
- Why is a negative test not always conclusive?
- List three precautions for a practical test.



14. Addition Reactions of Alkenes

The carbon–carbon double bond in an alkene can react with small molecules to form products in which new atoms or groups attach to the carbon atoms. The double bond becomes a single bond as the reaction proceeds. These are addition reactions and are characteristic of unsaturated hydrocarbons.

Hydrogenation adds hydrogen to an alkene in the presence of a suitable catalyst, often nickel under controlled industrial conditions. Ethene reacts with hydrogen to form ethane: $C_2H_4 + H_2 \rightarrow C_2H_6$. Hydrogenation is used in industrial processing, including some transformations of oils, although food processing requires careful control of product composition.

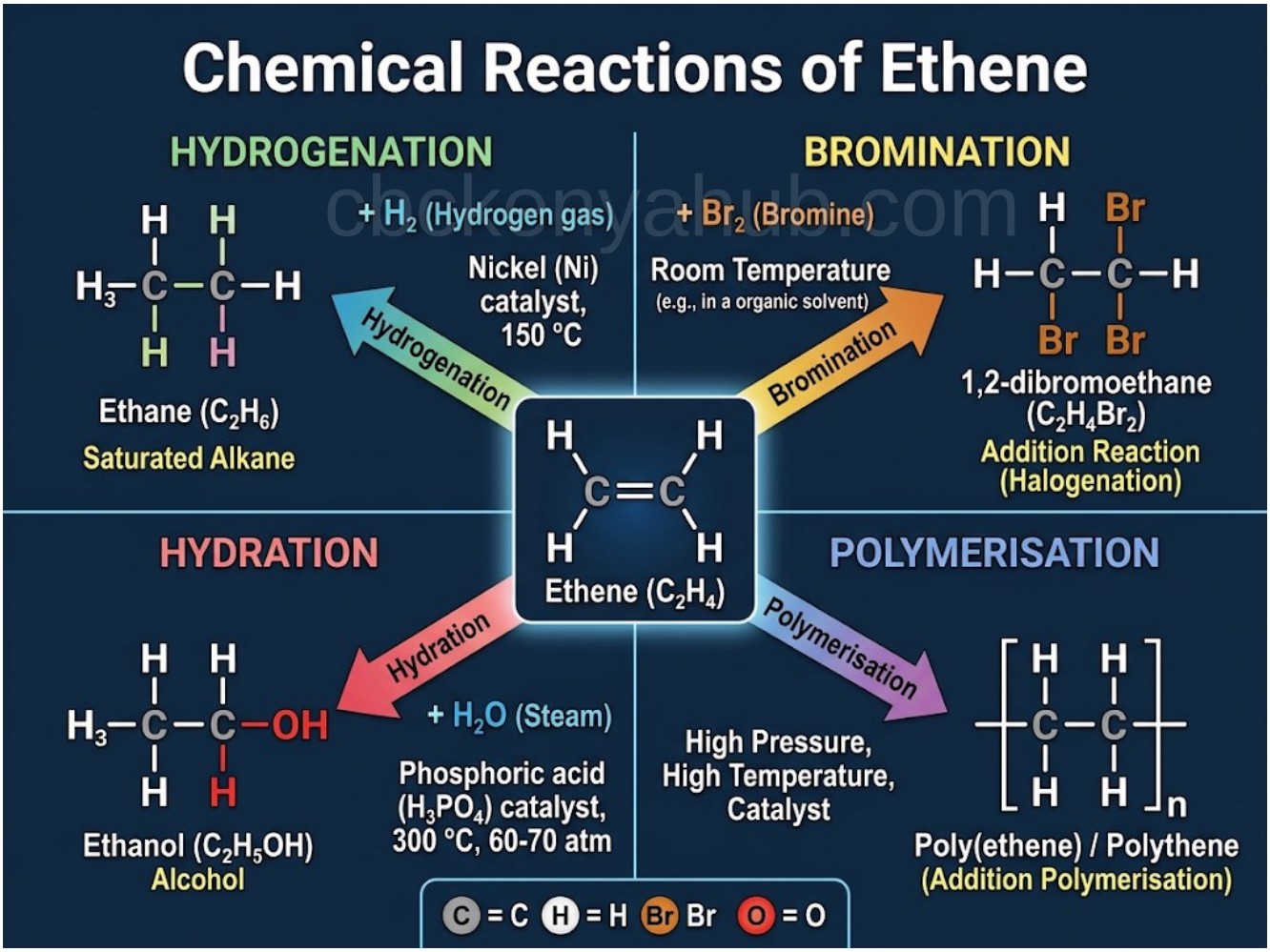
Halogen addition is illustrated by the reaction of ethene with bromine to form 1,2-dibromoethane. The colour change in bromine water is linked to the consumption of bromine during addition. The exact conditions depend on the reactants and medium.

Hydration adds water across the double bond to form an alcohol under suitable industrial conditions. Ethene can be converted to ethanol using steam and an acid catalyst at elevated temperature and pressure. The process is reversible and is managed to obtain a useful yield.

Addition polymerisation is another major reaction of alkenes. Many small alkene molecules (monomers) join to form a long-chain polymer. Ethene forms poly(ethene), widely used in packaging and many everyday objects. The polymer's properties depend on its chain structure and manufacturing conditions.

Check Your Understanding

- Define an addition reaction.
- Write the equation for hydrogenation of ethene.
- State the product of adding bromine to ethene.
- Explain the role of the double bond in addition reactions.



15. Alkynes: Structure and General Formula

Alkynes are unsaturated hydrocarbons containing at least one carbon–carbon triple bond. Ethyne, C_2H_2, is the simplest member and is commonly called acetylene. Its condensed formula is HC≡CH. The triple bond contains three shared electron pairs between the carbon atoms.

The general formula for a simple open-chain alkyne with one triple bond is C_nH_{2n-2}. For example, propyne is C_3H_4 and butyne is C_4H_6. This formula assumes an open chain with one triple bond and no additional rings or multiple bonds.

Because the triple bond needs two carbon atoms, an alkyne must contain at least two carbon atoms. With longer chains, the triple bond may occupy different positions and branches may occur, leading to structural isomers. The parent chain must include the triple bond when naming an alkyne.

Alkynes can undergo addition reactions because the multiple bond provides a site for reaction. Ethyne has been used as a fuel in oxy-acetylene welding and cutting because it can produce a very hot flame when burned with oxygen. This use requires purpose-designed equipment and strict safety procedures.

Do not attempt to generate, collect or ignite ethyne without formal laboratory supervision. Gas mixtures can be explosive, and pressurised gas cylinders require specialist handling. In this strand, focus on structure, formulae, naming and reaction equations rather than unsupervised gas preparation.

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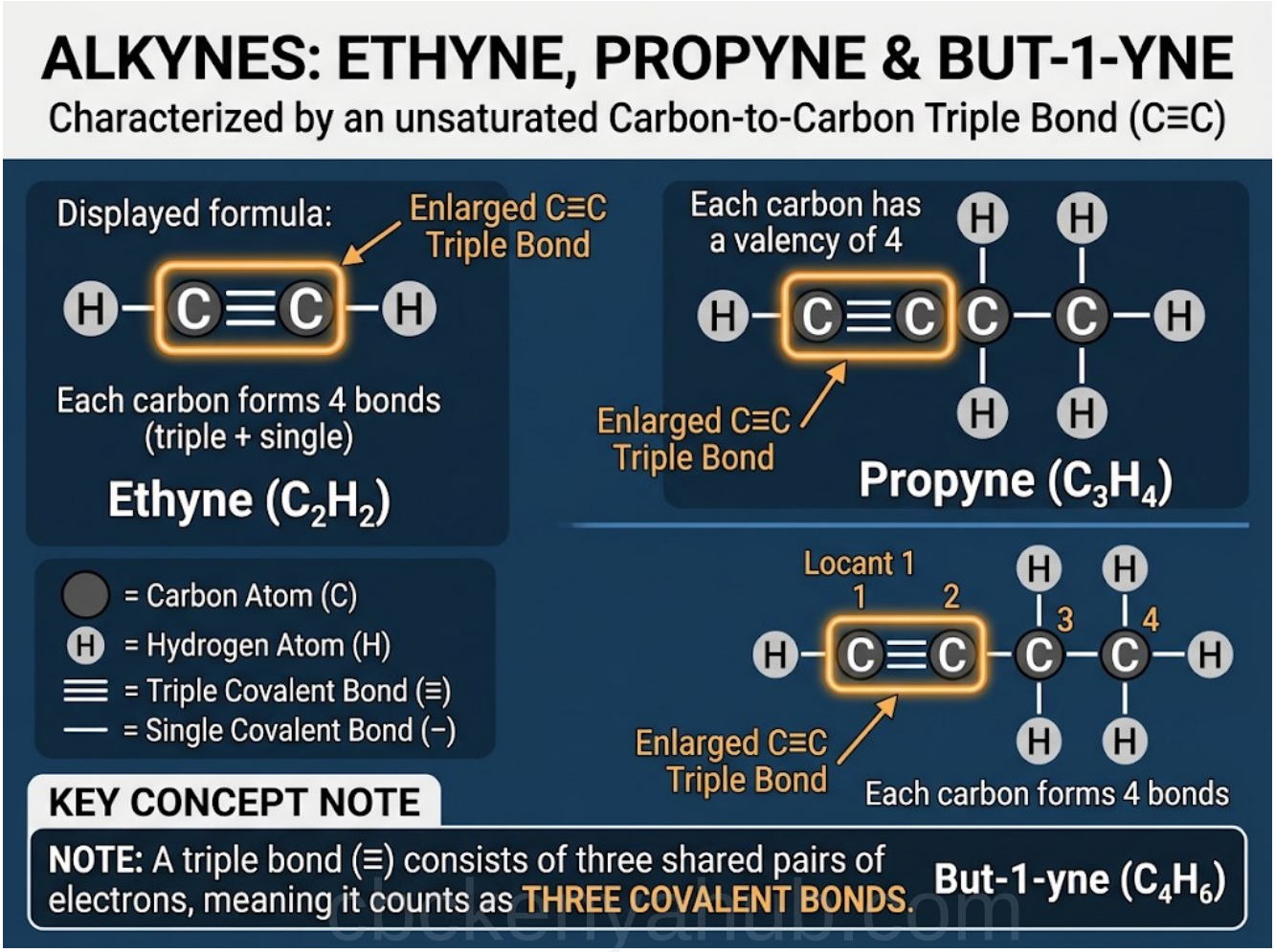
Check Your Understanding

- Define an alkyne.
- Write the general formula for an open-chain alkyne with one triple bond.
- Give the formula of ethyne and propyne.
- State one use of ethyne and one associated safety concern.



16. Naming Alkynes

Alkynes are named using the same basic principles as alkenes, but the ending is -yne. Choose the longest continuous chain containing the triple bond, then number from the end nearest the triple bond. Indicate the position of the triple bond using the number of the first carbon atom in it.

Examples include ethyne, propyne, but-1-yne and but-2-yne. In but-1-yne, the triple bond is between carbon 1 and carbon 2. In but-2-yne, it is between carbon 2 and carbon 3. The number must be as low as possible while the parent chain contains the triple bond.

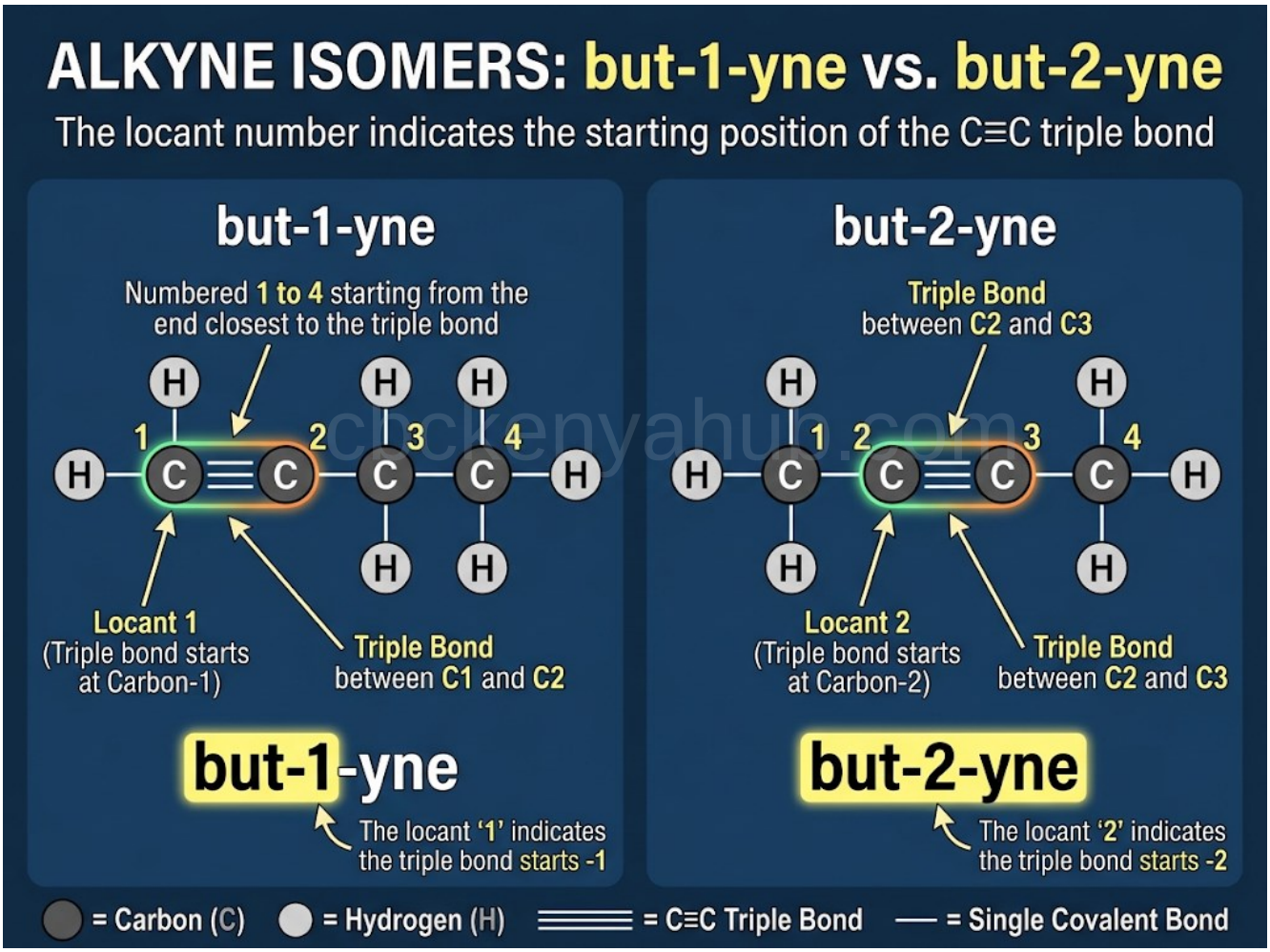
Branches are named and numbered using the same approach used for alkanes and alkenes. When a branched structure is given, first identify the longest chain that contains the triple bond. Numbering must prioritise the triple bond location over branch locations.

Worked example: Name $CH_3–C≡C–CH_3$. The parent chain contains four carbon atoms, so it is a butyne. The triple bond begins at carbon 2 from either end. The correct name is but-2-yne.

Check the chain length, triple-bond position, branch locants and suffix. A common error is naming a compound as an alkane because the learner counts only the carbon atoms and overlooks the triple bond. Another error is using -ene for a triple bond.

Check Your Understanding

- Name $HC{\equiv}C{-}CH_2{-}CH_3$.
- Name $CH_3{-}C{\equiv}C{-}CH_3$.
- Explain how the parent chain is selected for an alkyne.
- State the suffix used for alkynes.



17. Comparing Alkanes, Alkenes and Alkynes

Alkanes, alkenes and alkynes all contain carbon and hydrogen, but their bonding differs. Alkanes contain only single bonds and are saturated. Alkenes contain a carbon–carbon double bond and alkynes contain a carbon–carbon triple bond; both are unsaturated.

Their simple open-chain general formulae provide a useful check: alkanes C_nH_{2n+2}, alkenes with one double bond C_nH_{2n}, and alkynes with one triple bond C_nH_{2n-2}. For a given number of carbon atoms, the simple alkene has two fewer hydrogen atoms than the corresponding alkane; the simple alkyne has four fewer.

Alkanes commonly undergo combustion and substitution reactions. Alkenes and alkynes can undergo addition reactions at their multiple bonds. All three families can burn in oxygen, and all require careful handling because many members are flammable.

The bromine-water test can help identify unsaturation under appropriate conditions: an alkene typically decolourises bromine water, whereas an alkane does not rapidly do so in the absence of suitable reaction conditions. Classroom tests should be interpreted using observations and controls rather than assumptions.

Classification should be based on structure, not just on molecular formula. Formulae can be shared by different structural types, and rings or additional multiple bonds can change the expected hydrogen count. Use the formula as a check, then confirm the structure.

Check Your Understanding

- Compare the bonding in the three families.
- Calculate the molecular formula for hexane, hexene and hexyne under the stated assumptions.
- Explain the bromine-water test.
- Give one reaction characteristic of each family.

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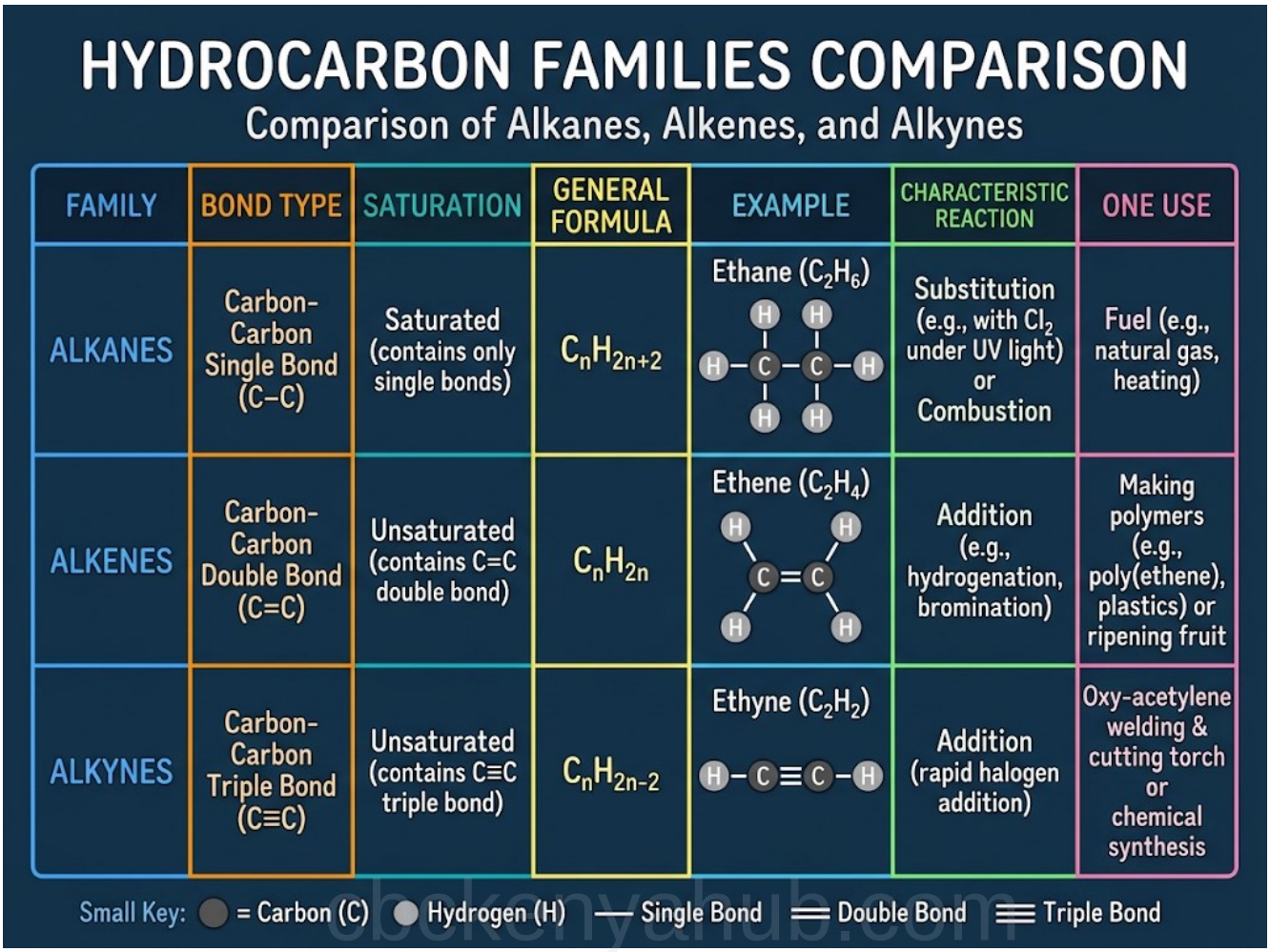


18. Petroleum and Natural Gas

Petroleum (crude oil) is a naturally occurring mixture of many hydrocarbons, formed over geological time from organic matter. Natural gas consists mainly of methane, although its exact composition varies by source and may include ethane, propane and other gases. These resources are important sources of energy and chemical feedstocks.

Crude oil is separated into fractions by fractional distillation. The mixture is heated so that much of it vaporises, and vapours rise through a fractionating column with a temperature gradient. Components condense at different levels according to their boiling-point ranges. The fractions are mixtures, not usually single pure compounds.

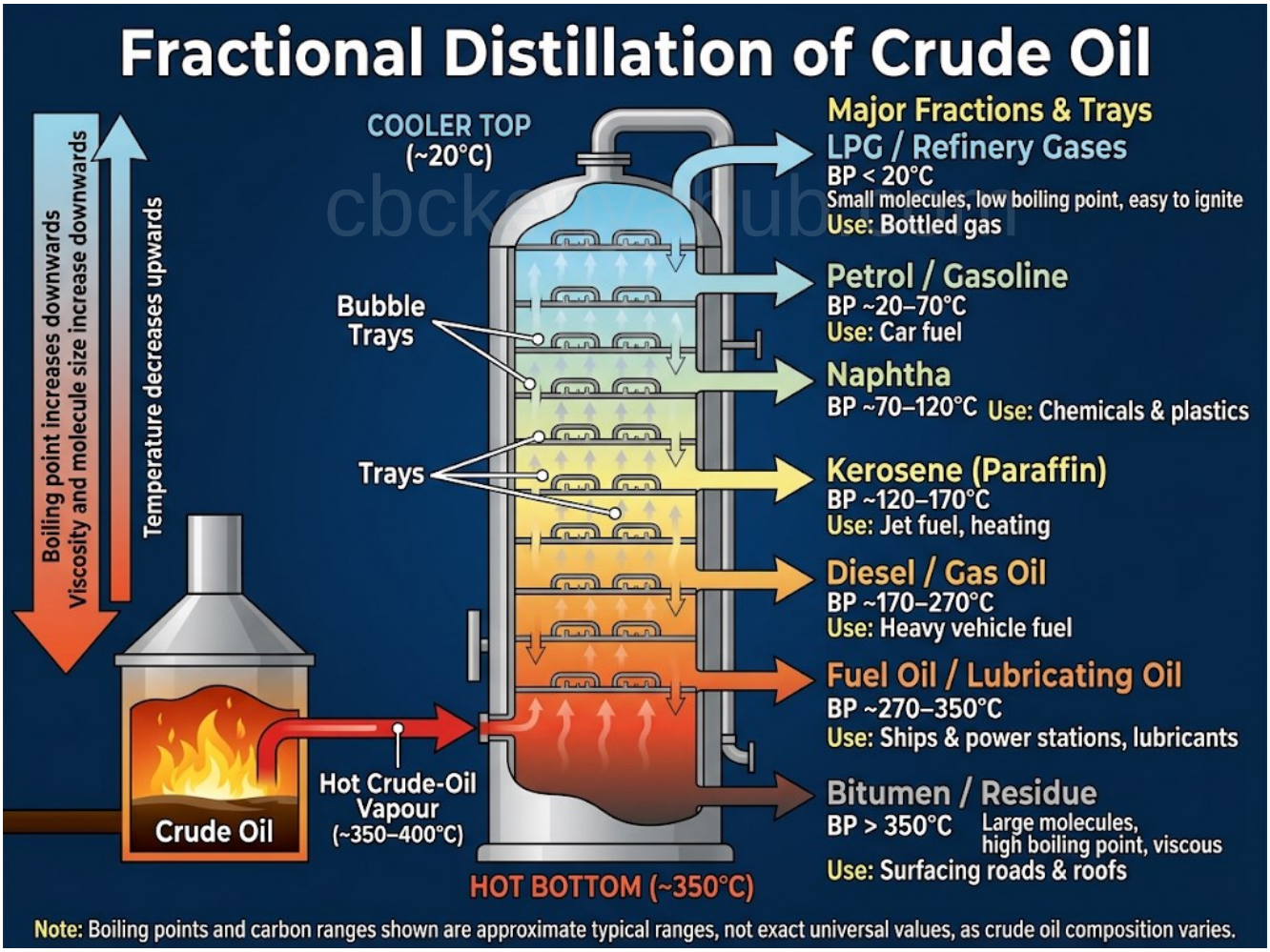
Smaller hydrocarbon molecules generally have lower boiling points and are more volatile than larger molecules in a homologous series. Refinery fractions may include refinery gases, petrol/gasoline, naphtha, kerosene, diesel, lubricating oils, fuel oils and bitumen. The exact fraction names and uses depend on the refinery and classification system.

Cracking breaks larger hydrocarbon molecules into smaller, more useful molecules, including alkanes and alkenes. It can increase the supply of lighter fuels and produce alkenes for polymer manufacture. Cracking requires suitable conditions and catalysts or high temperatures, depending on the process.

Fossil fuels are finite on human timescales and their combustion contributes to carbon dioxide emissions. Extraction and transport may also cause environmental damage. Energy efficiency, cleaner technology, responsible consumption and development of lower-carbon energy sources can reduce some impacts.

Check Your Understanding

- Explain why crude oil is separated by fractional distillation.
- Describe the temperature gradient in a fractionating column.
- Define cracking.
- State two benefits and two environmental concerns associated with fossil fuels.



19. Hydrocarbons as Fuels

Hydrocarbons release chemical energy when they react with oxygen. The energy released can be used for cooking, heating, transport and electricity generation. Different fuels have different energy densities, ignition properties, storage requirements, costs and emissions.

Complete combustion produces carbon dioxide and water, while incomplete combustion may produce carbon monoxide and soot. The likelihood of incomplete combustion increases when oxygen supply or mixing is inadequate. A clean-looking flame does not guarantee that all combustion products are harmless; carbon monoxide cannot be detected reliably by sight or smell.

Fuel quality is influenced by composition. Straight-chain and branched hydrocarbons may have different ignition characteristics in engines. Refining and blending are used to meet fuel specifications. The suitability of a fuel depends on the device and the required performance, not simply on whether it burns.

Environmental impacts include greenhouse gas emissions, local air pollutants and the risk of spills. Burning fuels in poorly ventilated rooms can be dangerous. Maintain appliances, use approved fuel containers, follow manufacturer guidance and ensure proper ventilation.

Energy comparisons should distinguish energy per unit mass from energy per unit volume. A fuel with high energy per kilogram may still require a large storage volume. In practical decisions, safety, availability, emissions and the technology using the fuel also matter.

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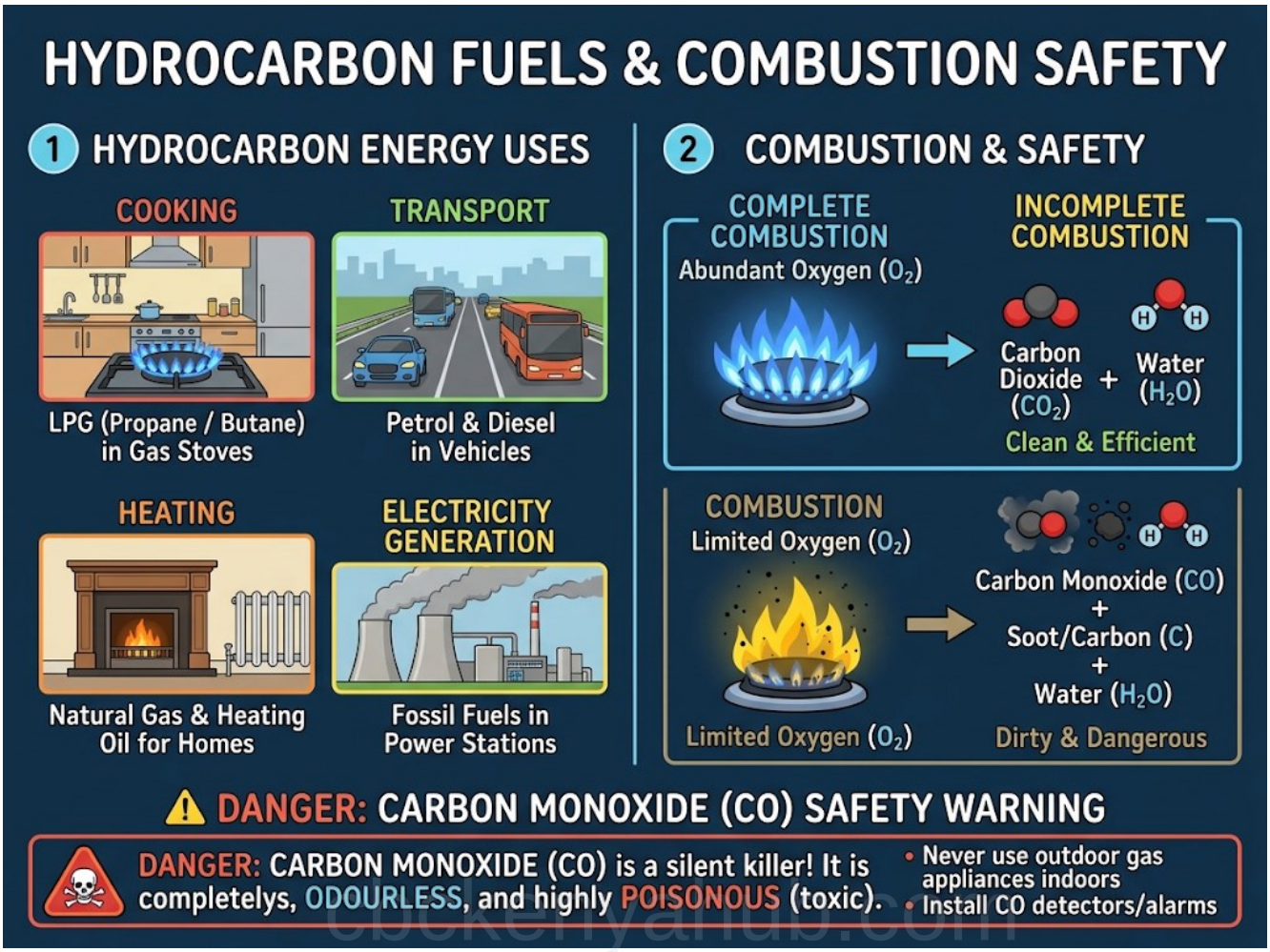
Check Your Understanding

- Explain why hydrocarbons are useful fuels.
- Distinguish complete and incomplete combustion products.
- Why is carbon monoxide difficult to detect?
- List four factors considered when choosing a fuel.



20. Addition Polymerisation

Addition polymerisation occurs when many unsaturated monomer molecules join to form a long-chain polymer. In alkene polymerisation, the carbon–carbon double bond opens and forms new carbon–carbon single bonds linking monomer units. No small molecule is eliminated in the ideal addition-polymerisation equation.

Ethene molecules form poly(ethene), also called polyethylene. The repeating unit is $–CH_2–CH_2–$, and the polymer chain is represented by brackets around the repeating unit with n indicating many units: $nCH_2=CH_2 \rightarrow [–CH_2–CH_2–]_n$. The symbol n does not specify one exact chain length; polymer samples contain chains of different lengths.

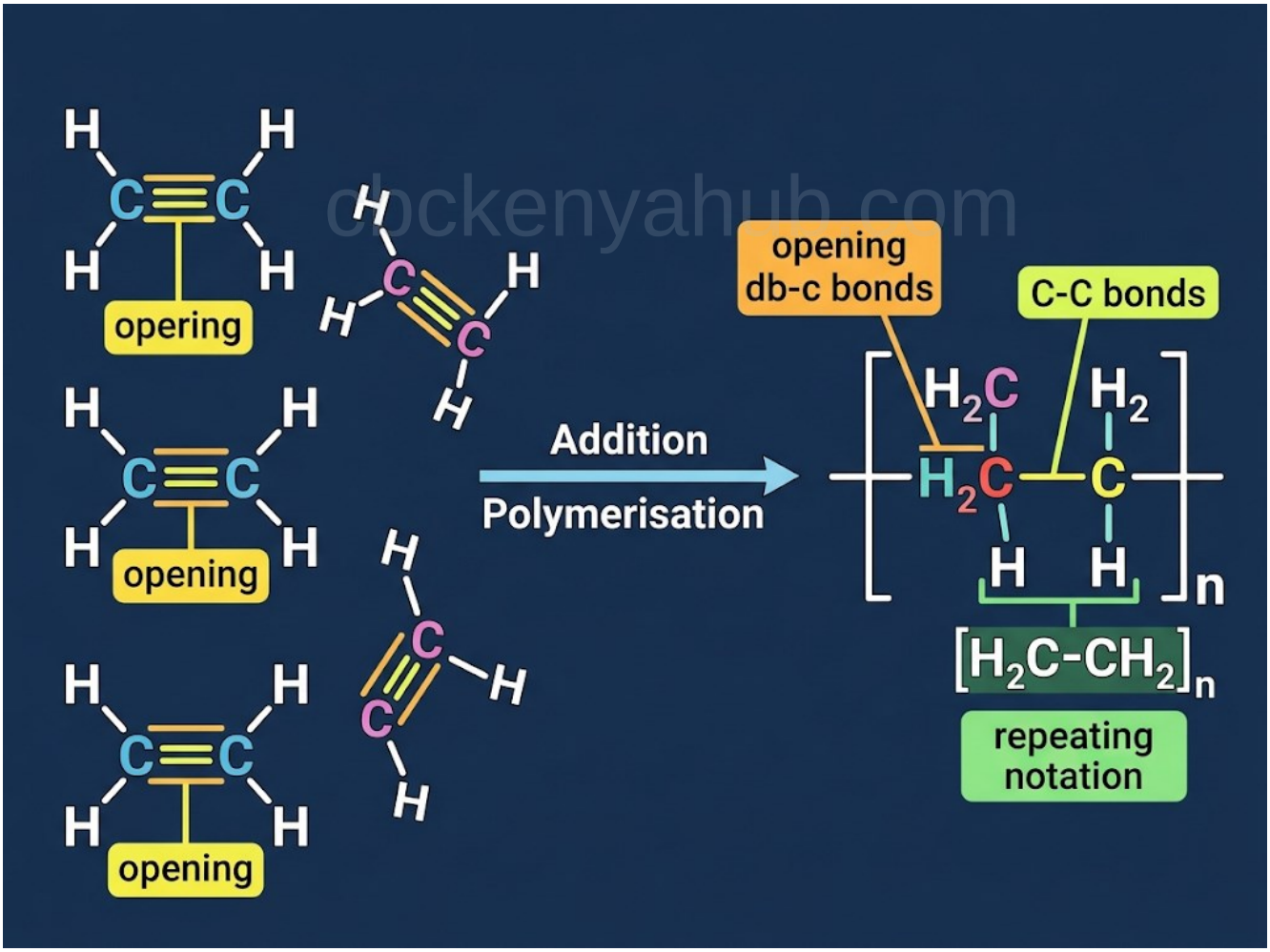
Propene forms poly(propene), whose repeating unit contains a methyl side group: $[–CH_2–CH(CH_3)–]_n$. The side group affects packing and properties. Different manufacturing conditions and polymer structures can produce materials with different stiffness, toughness, flexibility and melting behaviour.

Plastics are useful because they can be lightweight, durable, mouldable and resistant to corrosion. These same properties can create waste-management challenges. Some plastics persist in the environment, fragment into smaller particles and harm wildlife when mismanaged.

Responsible use includes reducing unnecessary single-use items, reusing products where safe, sorting waste correctly and supporting suitable recycling systems. Not every plastic can be recycled through every local programme; follow local collection guidance rather than assuming all plastic items belong in one recycling stream.

Check Your Understanding

- Define monomer and polymer.
- Write the polymerisation equation for ethene.
- Identify the repeating unit of poly(propene).
- Give three advantages and three environmental challenges of plastics.



21. Hydrocarbons in Everyday Life and Industry

Hydrocarbons are used directly as fuels and indirectly as raw materials. Refinery products provide fuels, lubricants, waxes, bitumen and feedstocks for chemical manufacture. Alkenes are especially important in producing polymers and other industrial chemicals.

Paraffin wax contains larger alkanes and is used in candles, coatings and some protective applications. Lubricating oils reduce friction between moving parts. Bitumen is used in road surfacing and roofing. These products are mixtures or materials whose properties depend on composition and processing.

Ethene is a key feedstock for poly(ethene), while propene is used to produce poly(propene). Hydrocarbon feedstocks can also be transformed into alcohols, solvents and other chemicals. In industrial chemistry, conversion steps are selected to achieve desired products while controlling cost, energy use, waste and safety.

Hydrocarbon solvents can dissolve oils, grease and other non-polar substances, but they may be volatile, flammable or harmful when inhaled or absorbed through skin. Always check product labels and safety instructions. Never use an unknown solvent on skin or near a flame.

Many products marketed as petroleum-derived are complex mixtures. Their use should follow the intended purpose and manufacturer instructions. The fact that a substance is derived from a natural source does not automatically make it harmless or environmentally benign.

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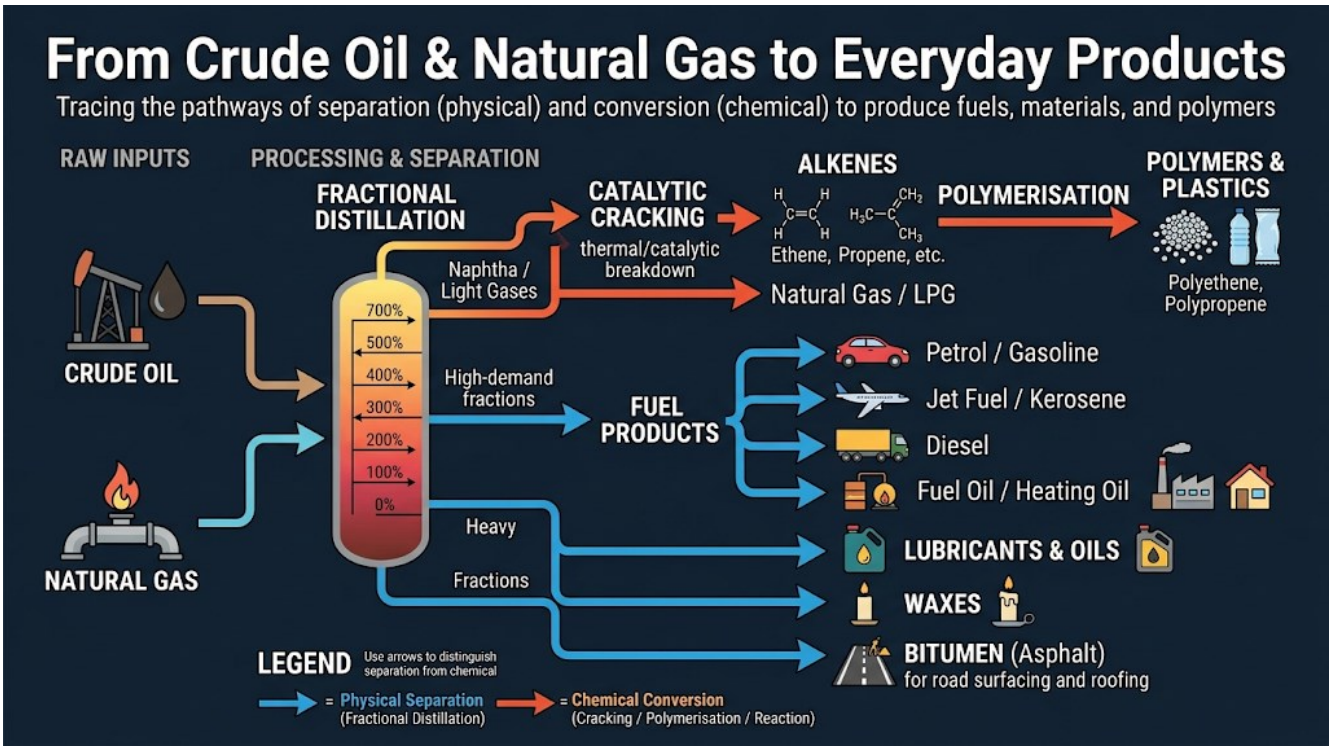
Check Your Understanding

- Name five everyday or industrial uses of hydrocarbons.
- Explain why alkenes are useful chemical feedstocks.
- Why should hydrocarbon solvents be handled carefully?
- Explain why natural origin does not guarantee safety.



22. Environmental Effects and Sustainability

Hydrocarbon use has supported transport, manufacturing, heating and electricity generation, but extraction, processing and combustion have environmental consequences. Carbon dioxide released during combustion contributes to the enhanced greenhouse effect. Incomplete combustion can release carbon monoxide and particulate matter, while some fuels and processes emit other pollutants.

Oil spills can coat shorelines, damage habitats and harm birds, fish and other organisms. Spilled hydrocarbons may spread over water, evaporate, dissolve partially or be broken down by microorganisms, depending on the compounds and conditions. Clean-up is complex and prevention is preferable to remediation.

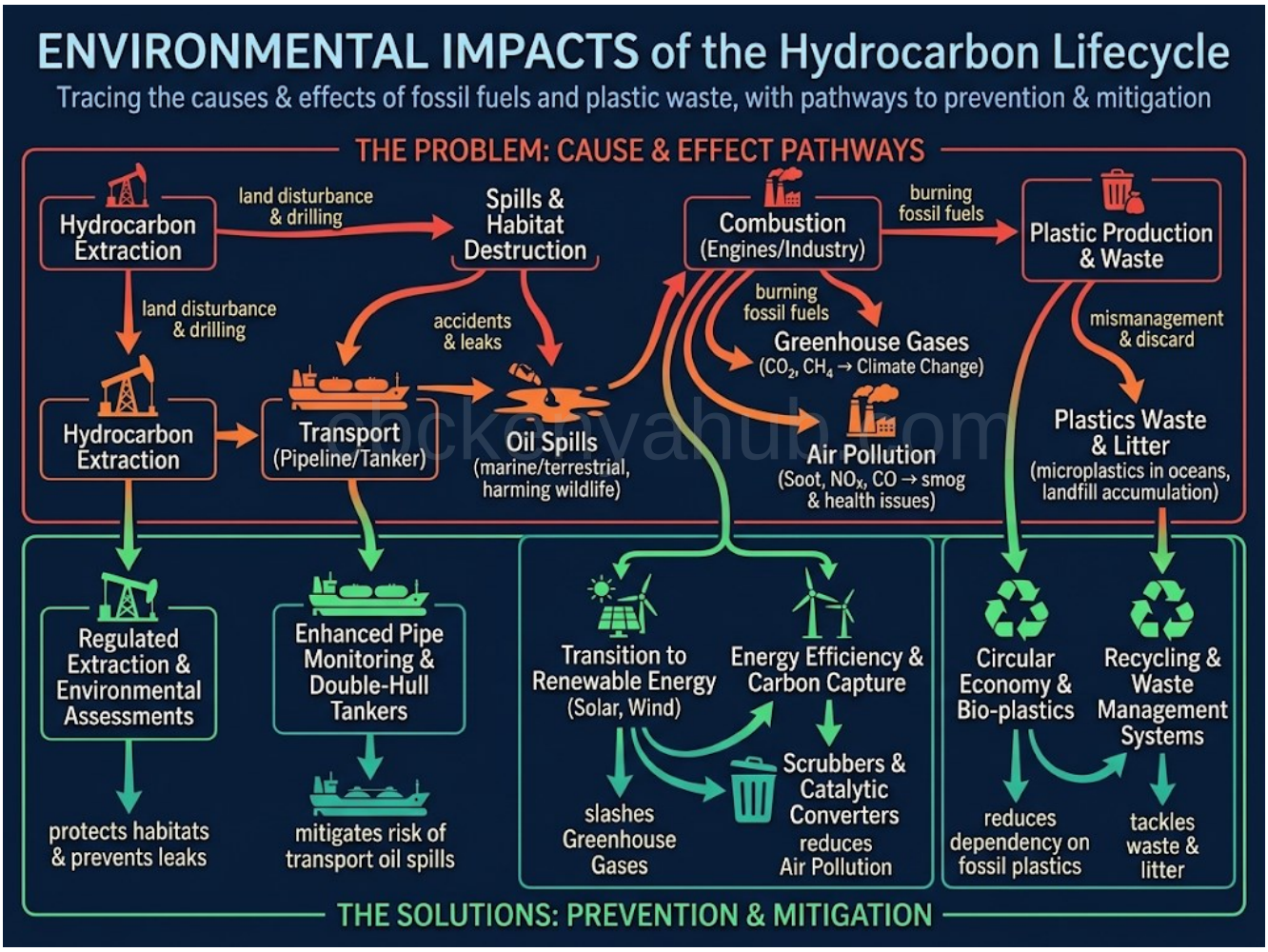
Natural gas can produce less carbon dioxide per unit of useful energy than some other fossil fuels when burned efficiently, but methane leakage matters because methane is a potent greenhouse gas over relevant time horizons. The full environmental comparison depends on the entire supply chain and the alternative energy source.

Sustainability measures include improving efficiency, preventing leaks, reducing waste, using public transport where practical, electrifying suitable applications and increasing the use of lower-carbon energy. Each measure has technical and social constraints, so decisions should be based on evidence and local context.

Chemistry helps explain both the benefits and the risks of hydrocarbons. Responsible choices require considering useful properties, energy needs, safety, environmental impacts and the feasibility of alternatives. A balanced discussion should avoid claiming that one fuel or technology has no impact.

Check Your Understanding

- Identify three environmental effects of hydrocarbon use.
- Explain why methane leakage matters.
- Give four measures that can reduce hydrocarbon-related impacts.
- Why is prevention of oil spills preferable to clean-up?



23. Laboratory Safety and Practical Skills

Practical chemistry requires risk assessment before work begins. Identify hazardous substances, likely exposure routes, ignition risks, glassware hazards and disposal requirements. Follow teacher

instructions and the approved method. Use eye protection and any additional protective equipment specified by the risk assessment.

Hydrocarbons are often flammable and volatile. Keep containers closed when not in use, use the smallest practical quantities, avoid inhaling vapours, and keep materials away from uncontrolled flames or sparks. Never heat a sealed container because pressure may rise dangerously.

When investigating combustion, use a stable apparatus and a controlled flame under supervision. Do not lean over a flame, tie back long hair, keep loose clothing away from heat and ensure flammable materials are cleared from the work area. If a fire occurs, alert the teacher and follow the school's emergency procedure.

For tests involving bromine water or other reagents, avoid skin and eye contact, use droppers carefully and dispose of waste as directed. Do not pour chemicals down a sink unless the approved procedure explicitly permits it. Clean spills only when instructed and equipped to do so.

A scientific report should separate observation from interpretation. For example, 'the orange-brown colour disappeared' is an observation; 'the sample contained an unsaturated compound' is a conclusion based on the test and its limitations. Include enough detail for another learner to understand what was done.

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Check Your Understanding

- List five steps in safe practical work.
- Why should hydrocarbon containers remain closed?
- Distinguish an observation from a conclusion.
- Explain why waste disposal must follow the approved method.

CLASSROOM CHEMISTRY SAFETY RULES



WEAR EYE PROTECTION

Always wear goggles to protect your eyes.

KEEP CONTAINERS CLOSED



Prevent spills and fumes by closing containers.

ENSURE GOOD VENTILATION



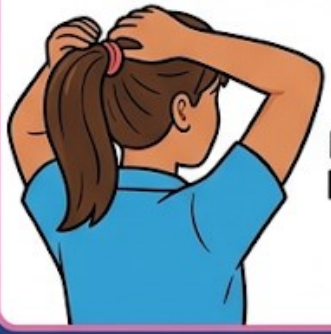
Work in well-ventilated areas or under fume hoods.

NO UNAUTHORIZED FLAMES



Never light a flame without the teacher's permission.

TIE BACK LONG HAIR



Keep hair and loose clothing tied back and secure.

CORRECT WASTE HANDLING



Dispose of chemical and solid waste in designated containers.

⚠ SAFETY WARNING: HYDROCARBON FLAMMABILITY

Hydrocarbons (like methane, propane, and ethanol) are highly flammable! Keep them far away from open flames, sparks, and heat sources.

24. Worked Examples: Formulae and Family Identification

Worked example 1: Determine the molecular formula of an open-chain alkane with six carbon atoms. The alkane formula is C_nH_{2n+2}. Substituting n = 6 gives C_6H_{14}. This compound is hexane if the chain is unbranched.

Worked example 2: Determine the formula of an open-chain alkene with one double bond and five carbon atoms. Use C_nH_{2n}. Substituting n = 5 gives C_5H_{10}. Several structural isomers may share this formula, so the formula alone does not identify one exact structure.

Worked example 3: Determine the formula of an open-chain alkyne with one triple bond and four carbon atoms. Use C_nH_{2n-2}. Substituting n = 4 gives C_4H_6. The position of the triple bond or branching may differ among isomers.

Worked example 4: Identify the family of C_3H_8. For n = 3, the alkane formula gives H = 2(3) + 2 = 8, so C_3H_8 fits propane, an alkane. Compare C_3H_6, which fits a simple alkene, and C_3H_4, which fits a simple alkyne.

Always state the assumptions used. The formula patterns given here apply to simple open-chain hydrocarbons with the specified number of multiple bonds. A cyclic compound or a compound with more than one multiple bond may have a different formula.

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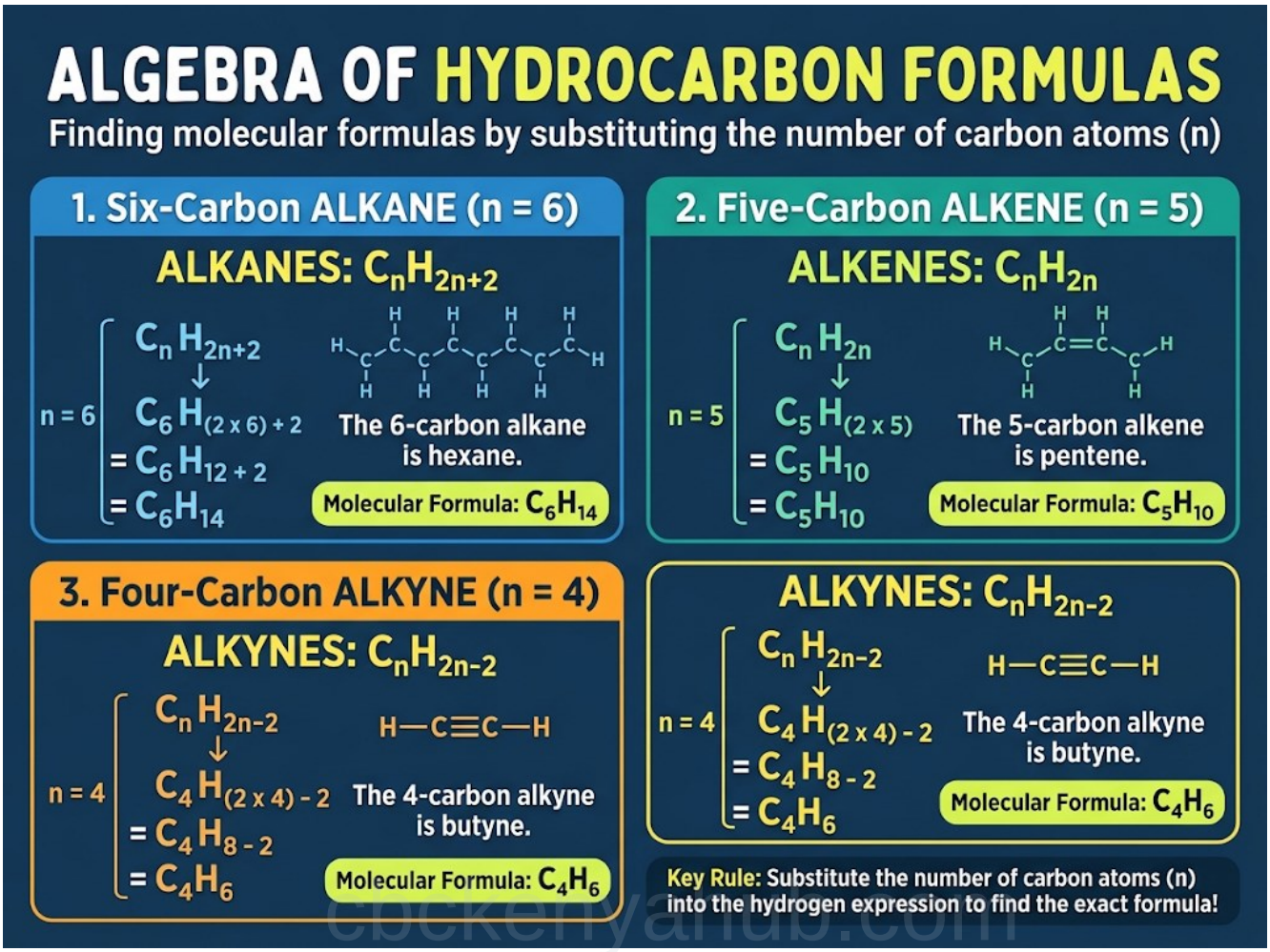
Check Your Understanding

- Calculate the formula of heptane.
- Calculate the formula of hexene with one double bond.
- Calculate the formula of pentyne with one triple bond.
- Explain why C_6H_{12} alone does not prove a sample is an alkene.



25. Worked Examples: Naming Structures

Worked example 1: $CH_3–CH_2–CH_2–CH_3$ contains four carbon atoms in a continuous chain and only single bonds. Its name is butane. There are no branches or locants to add.

Worked example 2: $CH_3–CH(CH_3)–CH_3$ has a three-carbon parent chain and a methyl branch on carbon 2. Numbering from either end gives the same locant. The name is 2-methylpropane.

Worked example 3: $CH_2=CH–CH_2–CH_3$ contains four carbon atoms in a chain that includes the double bond. Number from the end nearest the double bond, placing it at carbon 1. The name is but-1-ene.

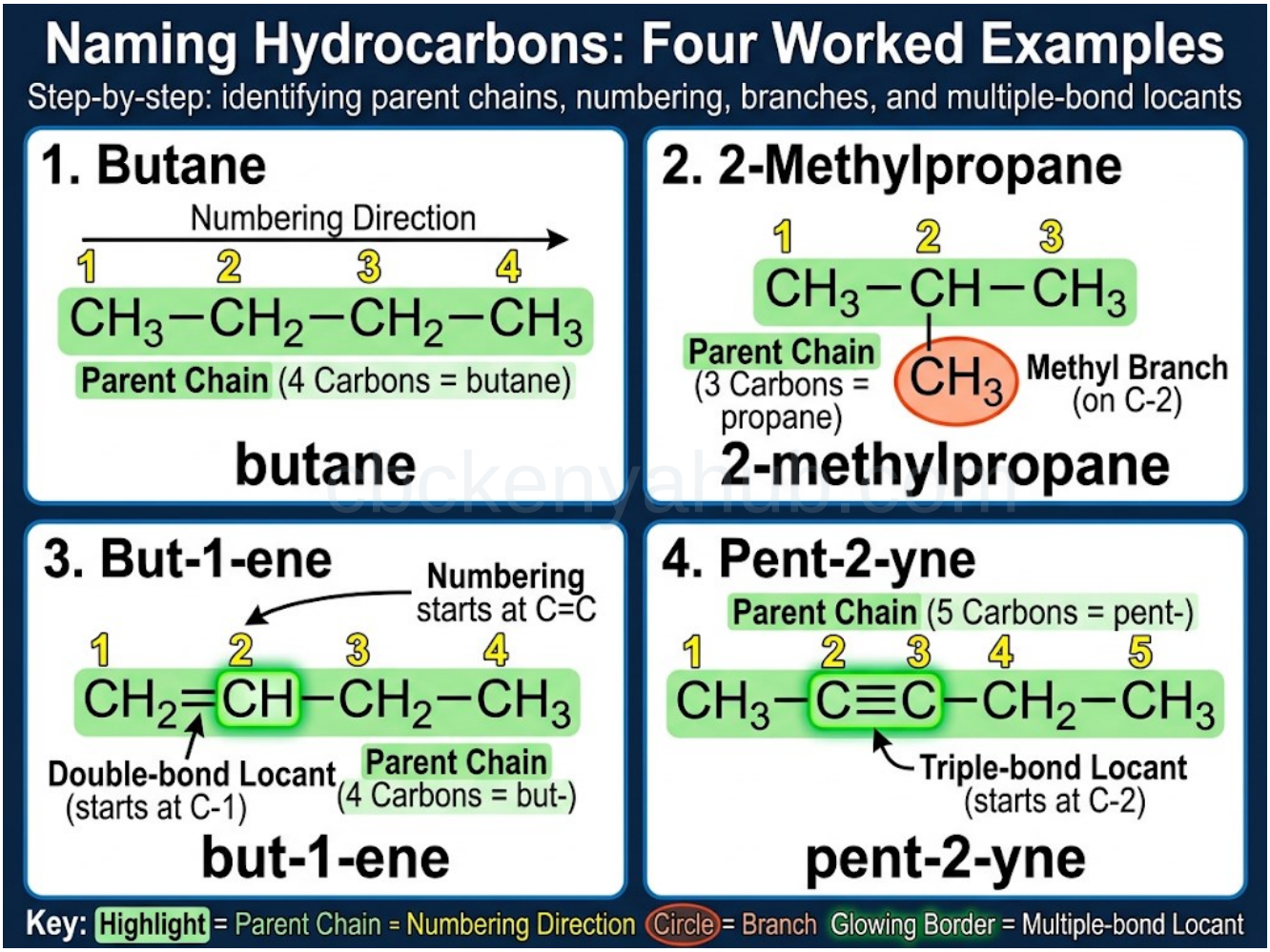
Worked example 4: $CH_3–C\equiv C–CH_2–CH_3$ has five carbon atoms and a triple bond beginning at carbon 2. The name is pent-2-yne. The triple bond must be included in the parent chain.

An efficient naming checklist is: (1) identify the family and multiple bond; (2) select the longest chain that contains the required bond; (3) number from the correct end; (4) locate and name branches; (5)

assemble the name with correct suffix, locants, commas and hyphens; (6) check the name against the structure.

Check Your Understanding

- Name $CH_3CH_2CH_2CH_2CH_3$.
- Name $CH_3CH(CH_3)CH_2CH_3$.
- Name $CH_3CH_2CH{=}CH_2$.
- Name $CH_3CH_2C{\equiv}CCH_3$.



26. Worked Examples: Balancing Equations

Combustion equations must conserve atoms. Begin by writing the correct products, then balance carbon, hydrogen and oxygen in that order. Do not change subscripts in chemical formulae to balance an equation; change only the coefficients placed in front of formulae.

Worked example: Complete combustion of butane. Start with $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$. Balance carbon: $C_4H_{10} + O_2 \rightarrow 4CO_2 + H_2O$. Balance hydrogen: $C_4H_{10} + O_2 \rightarrow 4CO_2 + 5H_2O$. There are 8 + 5 = 13 oxygen atoms on the product side, so the oxygen coefficient is 13/2 O_2. Multiply all coefficients by 2 to obtain $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$.

Worked example: Hydrogenation of propene. Propene is C_3H_6. Adding one molecule of hydrogen gives $C_3H_6 + H_2 \rightarrow C_3H_8$. The atoms balance: three carbon atoms and eight hydrogen atoms on each side.

Worked example: Addition of bromine to ethene. $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$. The product is 1,2-dibromoethane. The equation represents addition across the carbon–carbon double bond.

After balancing, count each type of atom on both sides. Include state symbols only when known or requested. Reaction conditions may be written above or below the arrow, but they do not replace the need for a balanced equation.

Check Your Understanding

- Balance the complete combustion of propane.
- Balance the complete combustion of pentane.
- Write the hydrogenation equation for but-1-ene.
- Explain why coefficients, not subscripts, are changed during balancing.

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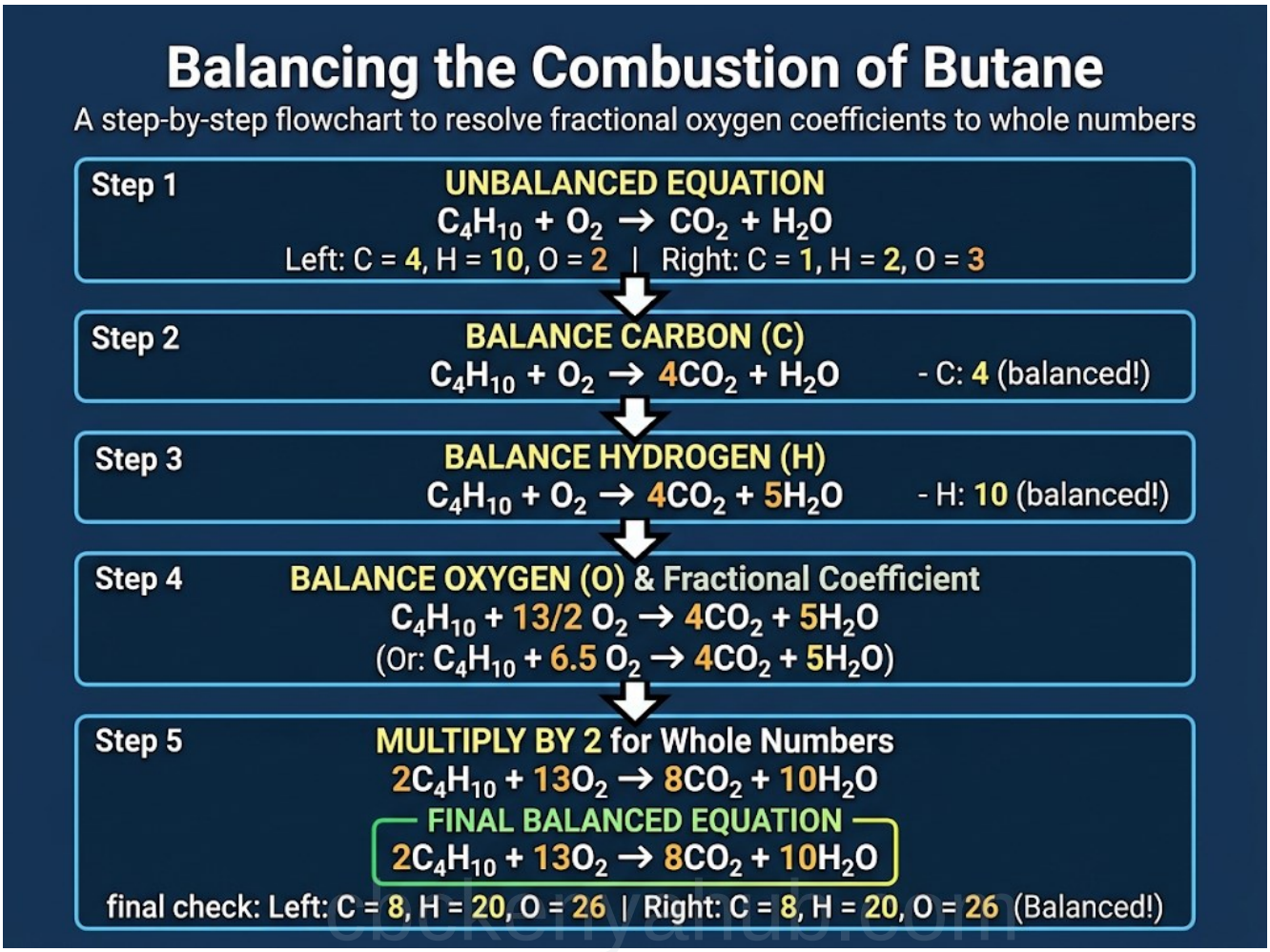


27. Revision Summary and Key Terms

Hydrocarbons contain only carbon and hydrogen. Alkanes are saturated and contain only single carbon–carbon bonds. Alkenes contain at least one carbon–carbon double bond. Alkynes contain at least one carbon–carbon triple bond.

The simple open-chain formulae are C_nH_{2n+2} for alkanes, C_nH_{2n} for alkenes with one double bond, and C_nH_{2n-2} for alkynes with one triple bond. These formulae are checks for specified structures, not universal formulae for every cyclic or multiply unsaturated hydrocarbon.

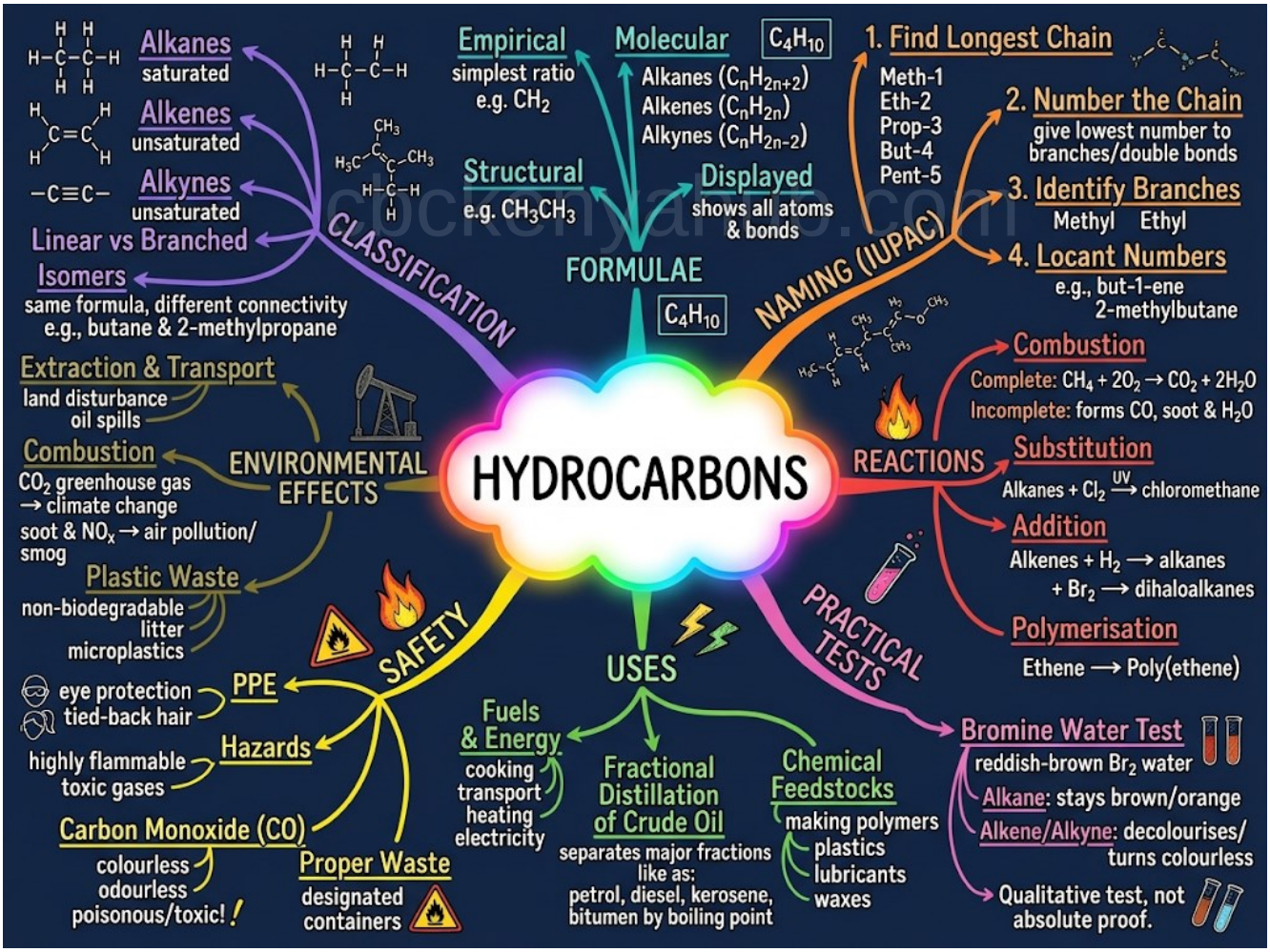
Naming depends on the longest appropriate parent chain, the position of multiple bonds, branch positions and the correct ending. Use -ane, -ene and -yne. Number the chain so that the multiple bond receives the lowest possible locant. Structural isomers share a molecular formula but differ in atom connectivity.

Complete combustion of a hydrocarbon forms carbon dioxide and water. Incomplete combustion can form carbon monoxide and soot. Alkanes can undergo substitution with halogens under suitable conditions, while alkenes and alkynes undergo addition reactions. Alkenes can also undergo addition polymerisation.

Hydrocarbons are valuable fuels and feedstocks, but their extraction, combustion and waste products create environmental and safety challenges. Scientific understanding supports safer handling, informed energy choices, waste reduction and responsible use of materials.

Check Your Understanding

- Define hydrocarbon, saturated, unsaturated, isomer, monomer and polymer.
- State the three general formulae and their assumptions.
- Compare substitution, addition and polymerisation.
- Give three safety precautions and three environmental concerns.



28. Revision Questions: Short Answer

1. Define a hydrocarbon and explain why carbon can form many different compounds. 2. Write the general formula for a simple open-chain alkane, alkene with one double bond and alkyne with one triple bond. 3. Name the first six straight-chain alkanes.

4. Draw displayed formulae for ethane, ethene and ethyne. 5. Explain the difference between a saturated and an unsaturated hydrocarbon. 6. Name $CH_3CH_2CH_2CH_3$ and $CH_3CH(CH_3)CH_3$. 7. Name $CH_2{=}CHCH_2CH_3$ and $CH_3C{\equiv}CCH_3$.

8. Define structural isomerism and draw the isomers of C_4H_{10}. 9. Explain why but-2-ene can have geometrical isomers but but-1-ene does not show the usual cis–trans isomerism. 10. Describe the bromine-water test for unsaturation.

11. Write balanced equations for complete combustion of methane and ethane. 12. Explain the difference between substitution and addition reactions. 13. Describe how crude oil is separated into fractions. 14. Define cracking and explain one industrial benefit.

15. Explain addition polymerisation using ethene. 16. State three uses of hydrocarbons and three risks associated with their use. 17. Explain why a fuel-burning appliance needs adequate ventilation. 18. Give two reasons why the bromine-water test must be interpreted cautiously.

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Check Your Understanding

- Answer each question in complete scientific sentences.
- Draw and label structures clearly.
- Show working for all formula and balancing questions.
- Use equations where a reaction is requested.

Name: _____ Date: _____

REVISION WORKSHEET: HYDROCARBON STRUCTURES

For each of the **hydrocarbons below:**

1. Draw the fully displayed formula in the large blank box.
2. Fill in the Labeled boxes below for its **Molecular Formula, Family,** and **Key Bond Type.**

1. ETHANE

DRAW DISPLAYED FORMULA HERE

Molecular Formula: [____]	Family: [_____]	Key Bond Type: [_____]

2. ETHENE

DRAW DISPLAYED FORMULA HERE

Molecular Formula: [____]	Family: [_____]	Key Bond Type: [_____]

3. ETHYNE

DRAW DISPLAYED FORMULA HERE

Molecular Formula: [____]	Family: [_____]	Key Bond Type: [_____]

4. BUTANE ISOMERS (Draw Both)

Butane (straight-chain)	**2-Methylpropane** (branched)

Molecular Formula: [____]	Family: [_____]	Key Bond Type: [_____]

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5. *cis*-BUT-2-ENE (Geometric Isomer)

DRAW DISPLAYED FORMULA HERE

Molecular Formula: [____]	Family: [_____]	Key Bond Type: [_____]

6. *trans*-BUT-2-ENE (Geometric Isomer)

DRAW DISPLAYED FORMULA HERE

Molecular Formula: [____]	Family: [_____]	Key Bond Type: [_____]

29. Revision Questions: Structured Problems

Question 1: A hydrocarbon has five carbon atoms and is an open-chain alkane. (a) Calculate its molecular formula. (b) Give its unbranched name. (c) Write a balanced equation for complete combustion. (d) Explain one danger of incomplete combustion.

Question 2: A compound has molecular formula C_4H_8 and decolourises bromine water. (a) Suggest the family it may belong to. (b) Draw two possible structural isomers. (c) Name one of the structures. (d) Explain why molecular formula alone may not identify one unique structure.

Question 3: An unknown compound has a carbon–carbon triple bond and four carbon atoms. (a) State its family. (b) Calculate the expected formula for an open-chain compound with one triple bond. (c) Draw but-1-yne and but-2-yne. (d) Explain the naming difference.

Question 4: Ethene is converted into a polymer. (a) Name the reaction type. (b) Draw the repeating unit. (c) Explain one useful property of the polymer. (d) Discuss one environmental challenge associated with plastic waste and one possible response.

Question 5: A learner tests two samples with bromine water. Sample A decolourises it quickly; sample B shows no rapid colour change under the same conditions. (a) State the likely distinction between the samples. (b) Explain the chemical basis of the test. (c) Identify one limitation. (d) State two precautions.

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Check Your Understanding

- Plan each answer before writing.
- Show calculations and balanced equations.
- Use labelled structural formulae rather than names alone.
- For discussion questions, give a reason for each point.

30. Answers and Self-Assessment

Suggested answers: Question 1: (a) C_5H_{12}. (b) Pentane. (c) $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$. (d) Incomplete combustion can produce poisonous carbon monoxide and soot. Question 2: (a) It may be an alkene with one double bond. (b) Possible examples include but-1-ene and but-2-ene; 2-methylpropene is also an isomer. (c) But-1-ene. (d) Different connectivities can share the same formula.

Suggested answers: Question 3: (a) Alkyne. (b) C_4H_6. (c) But-1-yne is $HC\equiv C–CH_2–CH_3$; but-2-yne is $CH_3–C\equiv C–CH_3$. (d) The locant identifies the first carbon in the triple bond. Question 4: (a) Addition

polymerisation. (b) [–CH_2–CH_2–]$_n$. (c) Poly(ethene) can be lightweight and mouldable. (d) Persistent waste can accumulate; reduction, reuse and appropriate recycling help.

Suggested answers: Question 5: (a) Sample A is consistent with an unsaturated compound such as an alkene; Sample B may be saturated under the test conditions. (b) Bromine adds across the carbon–carbon double bond, consuming coloured bromine. (c) Other reactive substances can also affect the reagent and a negative result is not conclusive by itself. (d) Use eye protection, small quantities and approved waste handling.

Self-assessment checklist: I can classify simple hydrocarbons; apply the general formulae with the correct assumptions; draw valid displayed and condensed formulae; name straight-chain and simple branched hydrocarbons; identify structural isomers; balance combustion equations; explain substitution, addition and polymerisation; describe the bromine-water test; and discuss safe and sustainable hydrocarbon use.

Where a learner is unsure, return to the relevant worked example and reproduce the method with a new compound. For naming, practise parent-chain selection before punctuation. For equations, count atoms on both sides. For practical questions, distinguish observations from conclusions and include relevant safety controls.

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Check Your Understanding

- Rate confidence from 1 (needs support) to 4 (can teach someone else).
- Choose two weak areas and revise them.
- Create three original naming questions and answer them.
- Explain one way chemistry can support safer and more sustainable use of hydrocarbons.

Glossary of Key Terms

Addition reaction: A reaction in which atoms or groups add to a molecule, often across a multiple bond.

Alkane: A saturated open-chain hydrocarbon with only carbon–carbon single bonds.

Alkene: An unsaturated hydrocarbon containing a carbon–carbon double bond.

Alkyne: An unsaturated hydrocarbon containing a carbon–carbon triple bond.

Branch: A carbon group attached to the parent carbon chain.

Combustion: Reaction with oxygen that releases energy.

Displayed formula: A representation showing atoms and covalent bonds.

Homologous series: A family of compounds with a common structural pattern and related properties.

Hydrocarbon: An organic compound containing only carbon and hydrogen.

Isomers: Compounds with the same molecular formula but different arrangements of atoms.

Monomer: A small molecule that can join with others to form a polymer.

Polymer: A large molecule made from many repeating monomer-derived units.

Saturated: Containing only carbon–carbon single bonds in the carbon skeleton.

Structural formula: A representation showing how atoms are connected.

Substitution: A reaction in which an atom or group is replaced by another atom or group.

Final Revision Checklist

☐ I can classify hydrocarbons by their bonds and apply the appropriate general formula.

☐ I can draw and interpret displayed and condensed formulae while respecting carbon valency.

☐ I can name straight-chain and simple branched alkanes, alkenes and alkynes.

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☐ I can explain structural isomerism and the conditions needed for cis–trans isomerism.

☐ I can balance combustion equations and describe selected addition reactions.

☐ I can explain the bromine-water test and identify its limitations.

☐ I can describe petroleum fractionation, cracking and addition polymerisation.

☐ I can discuss hydrocarbon uses, practical safety and environmental sustainability.