

Course Companion

International Baccalaureate · IB Diploma · Chemistry · HL

Every topic handout in one volume

exercises.lol

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Introduction to the particulate nature of matter

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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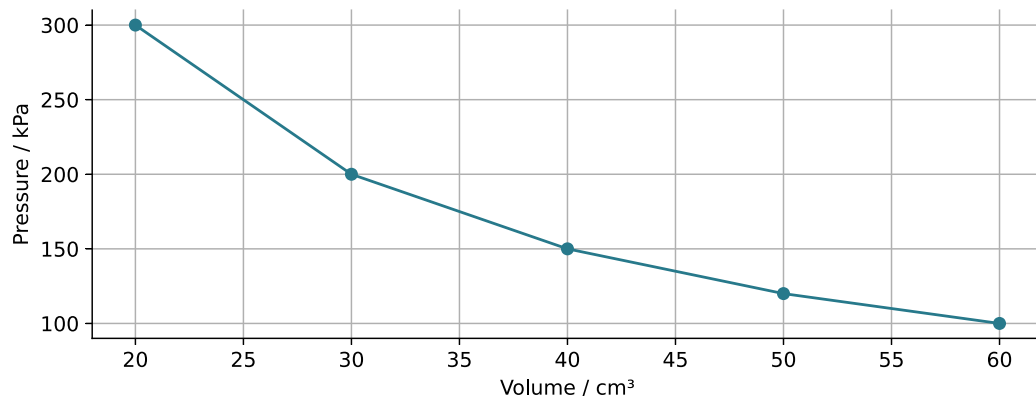
Gas models and absolute temperature

■ What would explain this observation?

- A sealed gas container changes pressure when heated. Celsius ratios cannot predict the pressure change because the gas model uses **absolute temperature**.

■ Build the model

- In a gas model, particles move randomly and pressure results from momentum transfer at walls.
- The **ideal gas** equation connects pressure, volume, amount and absolute temperature.
- absolute temperature: Temperature on the kelvin scale.
- ideal gas: A gas model with specified simplifying assumptions.



Gas models and absolute temperature: original worked-case diagram

■ Choose evidence that can test it

- At fixed amount and volume, pressure is proportional to kelvin temperature.
- At fixed temperature and amount, pressure is inversely proportional to volume.
- State which quantities are fixed before choosing a relationship.

In the lab: Use approved apparatus with a temperature range and pressure limit set by the teacher. Allow thermal equilibrium and record pressure against kelvin temperature. Never heat an improvised sealed vessel.

■ Work from known quantities

- Known: pressure is 100 kPa at 300 K, with fixed volume and amount.
- At 330 K, $p_2/p_1 = T_2/T_1$. $p_2 = p_1 \cdot T_2/T_1 = 100 \times 330/300 = 110$ kPa.
- A 30 °C rise is a 30 K change, but the temperature ratio must use kelvin.

Your turn: At fixed volume, pressure is 120 kPa at 300 K. Find pressure at 350 K.

■ Check the conclusion and its limits

- An ideal gas is a model with conditions of validity.
- Celsius zero is not zero molecular motion, and internal energy is not determined by pressure alone.

Not true: A gas pressure ratio can always use Celsius temperatures. Ratios of temperatures are meaningful only from absolute zero. 300 K to 350 K is a ratio of 350/300; the same temperatures in Celsius, 27 and 77, give 77/27 - a different, wrong answer.

Key:

- In a gas model, particles move randomly and pressure results from momentum transfer at walls.
- At fixed amount and volume, pressure is proportional to kelvin temperature.
- Watch for: an ideal gas is a model with conditions of validity.

■ Write from a pressure of 108 kPa

- At fixed volume, the pressure is 90 kPa at 250 K.
- Find the pressure at 300 K. The ratio must use kelvin.

The nuclear atom

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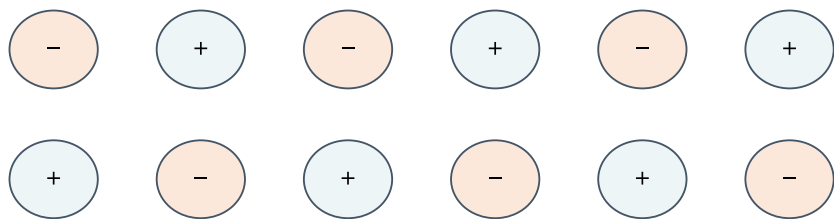
Particles, bonding and bulk properties

■ What would explain this observation?

- A salt crystal conducts when dissolved but not when solid. The ions exist in both states; their ability to move changes.

■ Build the model

- **Ionic bonding** is electrostatic attraction between oppositely charged ions.
- A covalent bond involves shared electrons.
- Metallic bonding involves attraction between positive metal ions and **delocalized electrons**.
- ionic bond : Attraction between oppositely charged ions.
- delocalized electron : An electron not confined to one atom or bond.



Alternating charges in an ionic lattice

Particles, bonding and bulk properties: original worked-case diagram

■ Choose evidence that can test it

- To explain a bulk property, name the structure, particles, forces and mobile charge carriers.
- Simple molecular substances can have strong covalent bonds inside molecules but weak attractions between molecules.

In the lab: Compare substances using evidence such as melting point, conductivity when solid and molten, and solubility. One property rarely proves a structure; use a pattern of evidence.

■ Work from known quantities

- Known: an element has atomic number 12 and mass number 24.
- Protons = 12; neutrons = mass number - atomic number = $24 - 12 = 12$.
- A $2+$ ion has electrons = $12 - 2 = 10$.
- Charge changes electron count, not the nucleus.

Your turn: An atom has atomic number 17 and mass number 35. Find its number of neutrons.

■ Check the conclusion and its limits

- Melting a simple molecular substance usually overcomes intermolecular attractions; it does not require breaking all covalent bonds within each molecule.

Not true: All covalent substances have low melting points. Giant covalent substances such as diamond and silicon dioxide have very high melting points because covalent bonds must be broken throughout the structure. Low melting points belong to simple molecular substances, where only intermolecular attractions are overcome.

Key:

- Ionic bonding is electrostatic attraction between oppositely charged ions.
- To explain a bulk property, name the structure, particles, forces and mobile charge carriers.
- Watch for: melting a simple molecular substance usually overcomes intermolecular attractions; it does not require breaking all covalent bonds within each molecule.

■ Write from 20 neutrons

- An atom has atomic number 19 and mass number 39.
- Find the neutron count and the electron count in the $1+$ ion. The charge does not change the nucleus.

Structure 1.3 · Electron configurations

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Scope and prerequisites

This handout covers selected HL/SL topics for the course first assessed in 2025; it is not the whole syllabus.

Prerequisites: atomic number and the structure of the atom from Structure 1.2, arithmetic with powers of ten, and the wave relation $c = f\lambda$.

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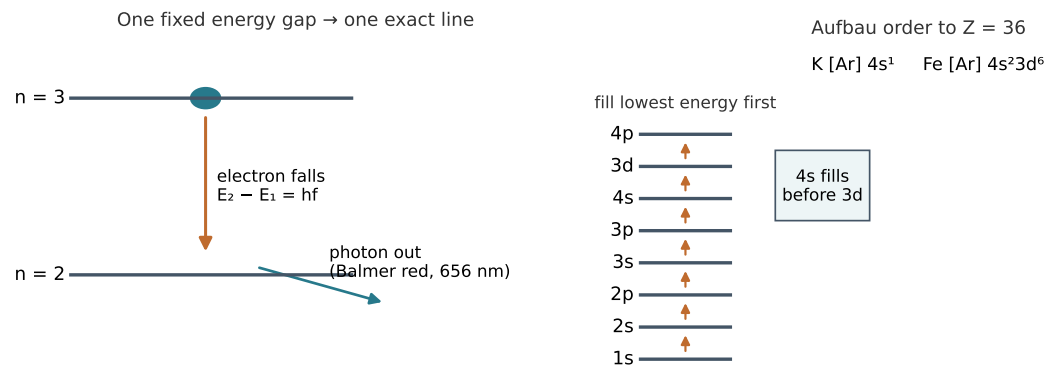
Electron configurations

■ What would explain this observation?

- A hydrogen discharge tube glows, and a prism splits the light into a few exact coloured lines – not a continuous rainbow. Why only those colours?

■ Build the model

- An electron in an atom can only occupy fixed **energy levels**. In a radiative transition to a lower level, it emits a photon whose energy equals the gap: $E_2 - E_1 = hf$. A line **emission spectrum** is the set of these exact photon energies, so each element shows its own pattern of lines.
- The usual filling-order guide is 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p. An **orbital** holds at most two electrons with opposite spins. The **electron configuration** lists the occupied sublevels, for example iron: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$, written compactly as $[\text{Ar}] 4s^2 3d^6$. Some ground-state configurations are exceptions to the simple filling guide; check the specified atom or ion.
- energy level: A fixed energy an electron may have in an atom; orbital: A spatial electron state holding at most two electrons with opposite spins; emission spectrum: The exact set of light frequencies an element emits.



One energy gap gives one spectral line; sublevels fill lowest energy first, 4s before 3d

■ Choose evidence that can test it

- The emission lines of hydrogen fall at exact frequencies predicted by the level gaps. If electrons could have any energy, the spectrum would be continuous – the observed separate lines are the evidence for fixed levels.
- The first ionisation energies of successive elements rise in steps and fall at the start of each new shell, matching the order in which sublevels fill. Compare a measured spectrum or ionisation pattern with the configuration model before accepting an explanation.

■ Work from known quantities

- Known: the red Balmer line of hydrogen has wavelength $\lambda = 656 \text{ nm} = 656 \times 10^{-9} \text{ m}$; $c = 3.00 \times 10^8 \text{ m/s}$ and $h = 6.63 \times 10^{-34} \text{ J s}$. The photon energy equals the level gap: $E = hf$ and $f = c/\lambda$, so $E = hc/\lambda$. Substitution: $E = (6.63 \times 10^{-34} \times 3.00 \times 10^8) / (656 \times 10^{-9}) = 3.03 \times 10^{-19} \text{ J}$. This is the energy gap between the $n = 3$ and $n = 2$ levels.

Example:

The ultraviolet line of hydrogen has $\lambda = 122 \text{ nm}$. Find the photon energy using the same sequence: known quantities → model → relation → substitution → unit and interpretation. (Answer: $E = 6.63 \times 10^{-34} \times 3.00 \times 10^8 / (122 \times 10^{-9}) = 1.63 \times 10^{-18} \text{ J}$ – about five times the red line's energy, because it spans the larger gap down to $n = 1$.)

■ Check the conclusion and its limits

- The filling order 4s before 3d is an empirical energy order for NEUTRAL atoms; when atoms lose electrons to form ions, the 4s electrons leave first (iron forms Fe^{2+} as $[\text{Ar}] 3d^6$, not $[\text{Ar}] 4s^2 3d^4$). The simple order also has noted exceptions such as chromium $[\text{Ar}] 4s^1 3d^5$.

Warn:

Electrons in an atom can occupy any energy between the shells. This claim is false: the separate lines of an emission spectrum appear only because electrons fall between fixed energy levels; a continuum of allowed energies would give a continuous spectrum, which is not observed.

Key:

Electron configurations: electrons occupy fixed energy levels, and the photon energy of each emission line equals the gap between two levels, $E = hc/\lambda$. Sublevels fill in the order 1s 2s 2p 3s 3p 4s 3d 4p, with at most two opposite-spin electrons per orbital.

Counting particles by mass: the mole

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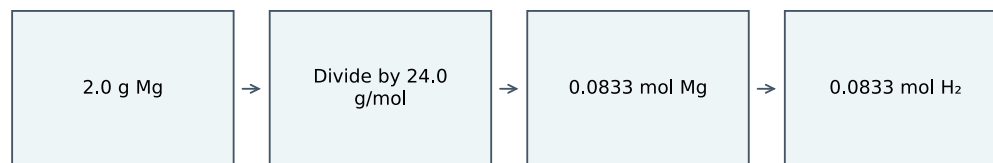
Amounts, equations and limiting reagents

■ What would explain this observation?

- The smallest mass of reactant is not necessarily the **limiting reagent**. The balanced equation compares particle amounts, not grams directly.

■ Build the model

- The **mole** measures amount of substance.
- Use molar mass to convert mass into amount.
- Balanced equation coefficients give mole ratios; they do not give equal masses.
- mole: The SI unit of amount of substance.
- limiting reagent: The reactant that limits the possible product amount.



Amounts, equations and limiting reagents: original worked-case diagram

■ Choose evidence that can test it

- Calculate the amount available for each reactant and divide by its coefficient.
- The smaller ratio limits the reaction.
- Use that reactant to calculate the maximum product before comparing actual yield.

In the lab: Write the balanced equation first, include units in molar masses, then convert each given mass or solution volume into amount. Convert cubic centimetres to cubic decimetres before using concentration in moles per cubic decimetre.

■ Work from known quantities

- Known: 2.0 g of Mg reacts with excess acid.
- Use amount = mass/molar mass.
- With molar mass Mg = 24.0 g per mole, amount Mg = $2.0/24.0 = 0.0833$ mol.
- In $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$, amount H_2 = amount Mg = 0.0833 mol.

Your turn: Calculate amount in 5.0 g of a substance with molar mass 100 g per mole.

■ Check the conclusion and its limits

- Excess acid means acid does not limit the stated calculation.
- A coefficient of 2 before HCl does not double the hydrogen amount.

Not true: Balanced equation coefficients always give reactant mass ratios. Coefficients give ratios of amounts in moles, not masses. Because each substance has a different molar mass, a 2:1 mole ratio is generally not a 2:1 mass ratio.

Key:

- The mole measures amount of substance.
- Calculate the amount available for each reactant and divide by its coefficient.
- Watch for: excess acid means acid does not limit the stated calculation.

■ Write from an amount of 0.30 mol

- A sample has mass 12.0 g and molar mass 40 g per mole.
- Find the amount. A mole ratio from an equation is not a mass ratio.

Ideal gases

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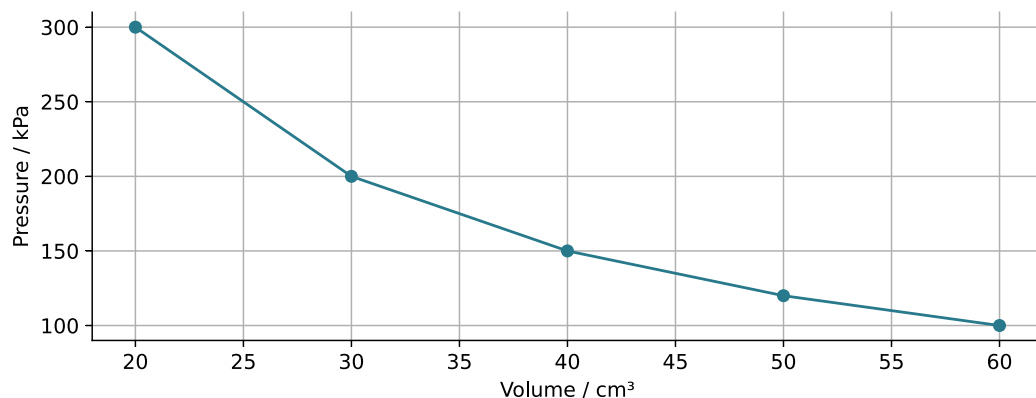
Ideal gases and the Kelvin scale

■ What would explain this observation?

- Doubling the kelvin temperature of a fixed amount of gas at fixed volume doubles its pressure. Celsius temperature shows no such simple proportionality. The **kelvin scale** starts at absolute zero.
-

■ Build the model

- **Ideal gas** behaviour links three quantities: pressure, volume and kelvin temperature.
- Temperature must be in kelvin: $K = \text{degrees Celsius} + 273$.
- At fixed amount and volume, pressure is proportional to kelvin temperature.
- At fixed temperature and amount, pressure is inversely proportional to volume.
- kelvin scale: The temperature scale starting at absolute zero; $K = \text{degrees Celsius} + 273$.
- pressure: Force per unit area exerted by gas particles colliding with the container walls.



Ideal gases and the Kelvin scale: Ideal gases: original worked-case diagram

■ Choose evidence that can test it

- State which quantities are fixed before choosing a relationship: proportionalities swap when the fixed quantities change.
- Record pressure against kelvin temperature with approved apparatus inside the teacher-set temperature range and pressure limit.
- Allow thermal equilibrium before each reading; never heat an improvised sealed vessel.

In the lab: Use approved apparatus with a temperature range and pressure limit set by the teacher. Allow thermal equilibrium and record pressure against kelvin temperature. Never heat an improvised sealed vessel.

■ Work from known quantities

- Known: a fixed amount of gas at 100 kPa and 300 K, volume held constant.
- At fixed volume and amount, pressure is proportional to kelvin temperature.
- Raise the temperature to 600 K: the ratio is $600/300 = 2$, so the pressure becomes 200 kPa.
- Cool instead to 150 K: the ratio is $150/300 = 0.5$, so the pressure becomes 50 kPa.

Your turn: A gas at 120 kPa and 300 K is cooled to 100 K at constant volume. Predict the new pressure.

■ Check the conclusion and its limits

- The proportionality $p \propto T$ holds only on the kelvin scale and only at fixed amount and volume.
- Real gases deviate from ideal behaviour at high pressure and low temperature; the model states its conditions.

Not true: Pressure is proportional to Celsius temperature. Proportionality holds on the kelvin scale only: 20 degrees Celsius is not twice as hot as 10 degrees Celsius, but 40 K is twice 20 K.

Key:

- Convert to kelvin before any proportionality.
- Name the fixed quantities, then choose the relationship.
- Watch for: real gases deviate at high pressure and low temperature.

■ Write from a volume of 12.3 dm³

- 0.500 mol of gas at 298 K and 101 kPa.
- $V = nRT$ over $P = 0.500 \times 8.31 \times 298 \div 101 = 12.3 \text{ dm}^3$ – R only works with kelvin and kPa.

■ Write from a doubled kelvin temperature

- A gas at 300 K in a rigid vessel is warmed to 600 K.
- The pressure doubles: P over T is constant at fixed volume – kelvin, never celsius, in the ratio.

■ Write from a deviation prediction

- The same gas is compressed to 5 000 kPa and cooled to 200 K.
- Does $PV = nRT$ hold? No – high pressure and low temperature push a real gas toward its condensed phase; the ideal model assumes warm and dilute.

The ionic model

International Baccalaureate · IB Diploma · Chemistry · HL

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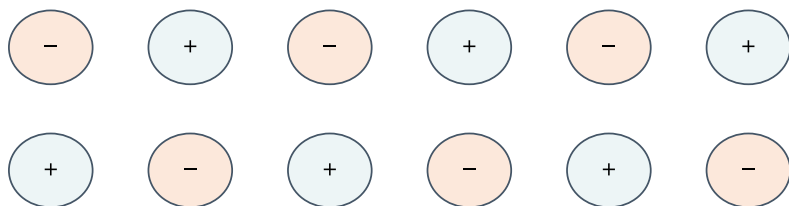
The ionic model: charged particles in a lattice

■ What would explain this observation?

- Solid sodium chloride does not conduct electricity; the molten salt does. The **ionic model** explains both facts with one structure.

■ Build the model

- Ions form by electron transfer: metal atoms lose electrons to non-metal atoms.
- Oppositely charged **ions** attract each other and pack into a giant **lattice**.
- In the solid the ions are fixed in place: no mobile charge carriers.
- When molten or dissolved, the ions move: charge carriers become mobile.
- ion: A charged particle formed by electron loss or gain.
- lattice: A regular repeating three-dimensional arrangement of particles.



Alternating charges in an ionic lattice

The ionic model: charged particles in a lattice: The ionic model: original worked-case diagram

■ Choose evidence that can test it

- To explain a bulk property, name the structure, the particles, the forces and the mobile charge carriers.
- Compare substances using melting point, conductivity when solid and when molten, and solubility.
- One property rarely proves a structure: use a pattern of evidence.

In the lab: Name structure, particles, forces and mobile charge carriers in that order, then match the pattern of melting point, conductivity (solid and molten) and solubility against the model.

■ Work from known quantities

- Known: substance X melts at 801 degrees Celsius.
- Solid X does not conduct; molten X conducts.
- Pattern: high melting point (strong electrostatic attraction in a lattice), no conduction when solid (fixed ions), conduction when molten (mobile ions).
- Conclusion: X fits the ionic model.

Your turn: Substance Y melts at 60 degrees Celsius and conducts neither solid nor molten. Which model fits Y better?

■ Check the conclusion and its limits

- The ionic model explains conduction and melting behaviour; solubility depends on the solvent as well as the structure.
- Simple molecular substances can have strong covalent bonds inside molecules but weak attractions between molecules.

Not true: A high melting point alone proves ionic bonding. A giant covalent substance also melts at a very high temperature: one property rarely proves a structure; use a pattern of evidence.

Key:

- Ions fixed in the solid, mobile when molten or dissolved.
- Explain bulk properties by naming structure, particles, forces and charge carriers.
- Watch for: one property alone proving a structure.

■ Write from a molten conductor

- Solid sodium chloride does not conduct; molten sodium chloride does.
- The charge carriers are the same ions – free to move only when the lattice melts: mobility, not existence, was the barrier.

■ Write from a melting-point comparison

- Sodium chloride melts at 801 C; magnesium oxide at 2 852 C.
- Mg²⁺ and O²⁻ carry 2+/2- charges against Na⁺/Cl⁻: the stronger electrostatic attraction holds the lattice to a far higher temperature.

■ Write from a one-property trap

- A white crystalline solid dissolves in water.
- Does that prove a giant ionic structure? It is consistent, not proof – solubility, melting point and conductivity TOGETHER identify the lattice.

The covalent model

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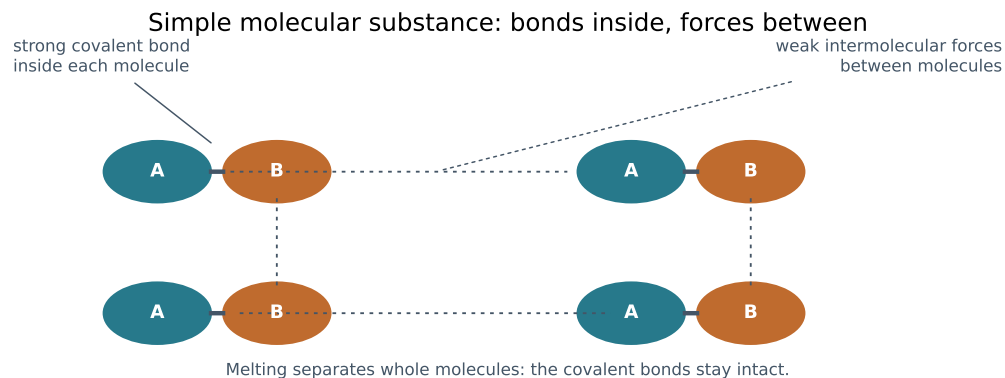
The covalent model: shared pairs between atoms

■ What would explain this observation?

- Iodine melts at 114 degrees Celsius even though its **covalent bonds** are strong. The bulk property is set by the forces BETWEEN molecules, not the bonds inside them.

■ Build the model

- A covalent bond is a shared pair of electrons between two nuclei.
- Simple molecular substances: strong covalent bonds inside each molecule.
- Between separate molecules the attractions are weak **intermolecular forces**.
- Melting and boiling separate molecules from each other, so bulk properties stay low.
- covalent bond: A shared pair of electrons holding two nuclei together.
- intermolecular forces: Weak attractions acting between separate molecules.



Simple molecular substance: bonds inside, forces between: original worked-case diagram

■ Choose evidence that can test it

- To explain a bulk property, name the structure, the particles, the forces and the mobile charge carriers.
- Simple molecular substances: low melting and boiling points, no conduction solid or molten (no ions, no mobile charge carriers).
- Solubility varies with polarity; one property rarely proves a structure.

In the lab: Distinguish the bonds **INSIDE** molecules from the forces **BETWEEN** them before explaining any melting or boiling point: strong bonds can coexist with a low boiling point.

■ Work from known quantities

- Known: substance A melts at minus 78 degrees Celsius; substance B melts at 114 degrees Celsius; both are simple molecular.
- Neither conducts when solid or molten: no mobile charge carriers in either.
- The difference: B's intermolecular forces are stronger, so more energy is needed to separate its molecules.
- The covalent bonds inside the molecules are not being broken at either melting point.

Your turn: Why does methane boil far below water though both are simple molecular? Name the forces involved.

■ Check the conclusion and its limits

- The covalent model here covers simple molecular substances; giant covalent networks behave differently.
- Intermolecular force strength varies between substances; the model names the difference, not a fixed value.

Not true: Strong covalent bonds give a substance a high melting point. Melting a simple molecular substance separates molecules from each other: the covalent bonds inside the molecules stay intact.

Key:

- Bonds inside molecules are strong; forces between molecules are weak.
- Melting and boiling act on the between-molecule forces.
- Watch for: explaining a molecular melting point by breaking covalent bonds.

■ Write from a boiling-point order

- Water boils at 100 C; hydrogen sulfide at minus 60 C.
- Both are bent molecules, but hydrogen bonding between O–H molecules dwarfs the dipole forces of H₂S – boiling acts **BETWEEN** molecules, never on the covalent bond within.

■ Write from a diamond-graphite contrast

- Both are giant covalent carbon, yet diamond is hard and non-conducting while graphite is soft and conducting.
- Diamond bonds each atom to four others; graphite layers carry three bonds plus delocalised electrons – same atoms, different framework.

■ Write from a state prediction

- A simple molecular substance with only weak London forces between its molecules.
- Predict a low melting point: the bonds **WITHIN** molecules are strong, but melting never breaks them – it separates molecules.

The metallic model

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The metallic model: cations in a sea of electrons

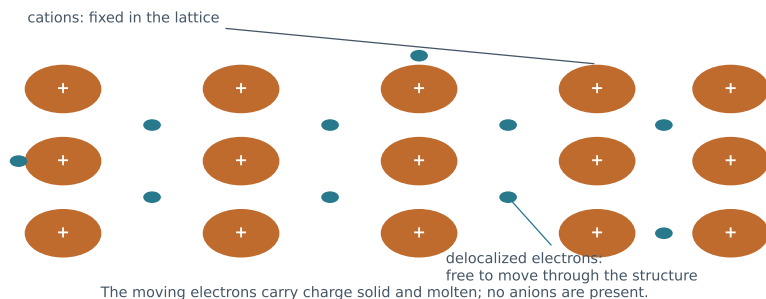
■ What would explain this observation?

- A copper wire conducts electricity and bends without shattering. The **metallic model** explains both with one picture: cations held by moving electrons.

■ Build the model

- Metal atoms lose outer electrons, leaving positive **cations** in a lattice.
- The lost electrons become **delocalized electrons**: free to move through the whole structure.
- The **metallic bond** is the attraction between the cation lattice and the delocalized electrons.
- The moving electrons carry charge in the solid and in the molten metal.
- cation: A positively charged ion formed by electron loss.
- delocalized electrons: Electrons free to move through the whole metallic structure.
- metallic bond: Attraction between a lattice of cations and delocalized electrons.

Metallic model: cation lattice, delocalized electrons



Metallic model: cation lattice, delocalized electrons: original worked-case diagram

■ Choose evidence that can test it

- To explain a bulk property, name the structure, the particles, the forces and the mobile charge carriers.
- Metals conduct when solid AND molten: the charge carriers are the delocalized electrons, mobile in both states.
- Malleability: bonding is non-directional, so layers can slide and the bonding reforms; ionic lattices shatter because sliding brings like charges together.

In the lab: Conduction in both solid and molten states distinguishes a metal from an ionic substance (which conducts only when molten or dissolved): one observation, two models separated.

■ Work from known quantities

- Known: substance Z melts at 1538 degrees Celsius and conducts both solid and molten.
- Charge carriers: delocalized electrons, mobile in both states.
- The metallic model predicts both observations with one structure.
- An ionic model would fail here: fixed ions in the solid cannot conduct.

Your turn: An ionic solid and a metallic solid look alike. State one measurement that separates them and explain it.

■ Check the conclusion and its limits

- Melting points vary widely between metals; the model explains conduction and malleability, not a fixed melting temperature.
- Alloys and impurities change properties: the pure model states its limits.

Not true: Metals conduct because electrons can leave the metal. The electrons are delocalized WITHIN the structure: conduction needs movement through the metal, not escape from it.

Key:

- Cations in a lattice, delocalized electrons moving through it.
- Conduction in solid and molten states marks the metallic model.
- Watch for: electrons needing to leave the metal to conduct.

■ Write from a conduction test

- A sample conducts as a solid AND as a liquid.
- Metallic: delocalised electrons carry the charge in both states – the model’s signature, no ions to release.

■ Write from a malleability explanation

- Copper bends; sodium chloride shatters.
- The cation layers slide while the electron glue keeps holding – metallic bonding explains malleability; a rigid lattice explains brittleness.

■ Write from the carrier contrast

- Molten NaCl conducts; solid NaCl does not; copper always does.
- The contrast is the carrier: ions freed by melting against electrons already free – one experiment separates ionic from metallic.

Structure 2.4 · From models to materials

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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Prerequisites: the ionic, covalent and metallic models from Structure 2.1–2.3, and density from earlier physical science.

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From models to materials

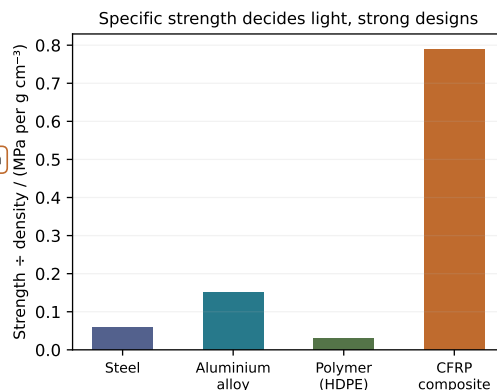
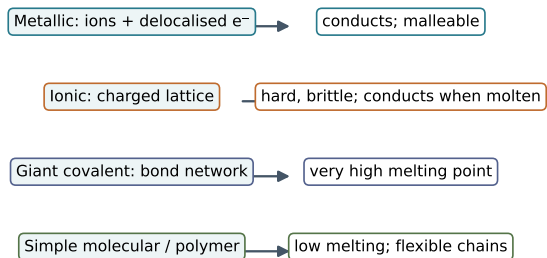
■ What would explain this observation?

- A copper wire bends without breaking and carries current; a ceramic plate snaps but never conducts; a plastic bag stretches and melts over a candle. Why do these materials behave so differently?

■ Build the model

- The bonding model predicts the **bulk properties** : a **metallic material** has positive ions in a sea of delocalised electrons, so it conducts and is malleable. An ionic lattice is hard but brittle and conducts only when molten or dissolved. A giant covalent network has a very high melting point. Simple molecular substances and polymers melt low, and polymer chains can slide and flex.
- A **composite** combines materials so the product has properties neither part has alone, such as carbon fibres in a resin. Choosing a material means matching its model-derived properties to a design requirement, then checking measurable quantities such as strength, density and conductivity.
- bulk property: A measurable property of a material sample, such as strength or conductivity; composite: A material made of two or more components whose combined properties exceed the parts; metallic material: A material of metal ions held by delocalised electrons.

The bonding model predicts the material's properties



From models to materials: original worked-case diagram

■ Choose evidence that can test it

- Test the model by measuring what it predicts: bend and stretch samples of a metal, an ionic solid and a polymer under the same conditions and compare conductivity in solid and molten states where the school permits. The model is supported only if the observed properties match the bonding in the structure.
- Compare candidate materials for one application using measured properties, not names: a bicycle frame needs high strength per unit mass, a window needs transparency and scratch resistance.

■ Work from known quantities

- Known: an aluminium alloy has strength 300 MPa and density 2.7 g/cm³; a steel has strength 500 MPa and density 7.8 g/cm³. The design needs strength per unit mass, so compute the ratio: alloy 300/2.7 $\approx$ 111 MPa per g/cm³; steel 500/7.8 $\approx$ 64 MPa per g/cm³. The alloy gives more strength for the same mass, so it suits a light frame even though the steel is stronger in absolute terms.

Example:

A carbon-fibre composite has strength 1,100 MPa at density 1.4 g/cm³. Find its strength-to-density ratio and compare it with the alloy above, using the same sequence: known quantities → model → relation → substitution → unit and interpretation. (Answer: 1,100/1.4 $\approx$ 786 MPa per g/cm³ – about seven times the alloy, which is why racing frames use it despite the cost.)

■ Check the conclusion and its limits

- A single ratio never settles a design choice: toughness, cost, corrosion resistance, manufacturability and end-of-life all matter. Bonding models describe ideal structures; real materials contain defects that lower measured strength below the model's prediction.

Warn:

The strongest material in absolute terms is always the best choice for a design. This claim is false: the best choice depends on the requirement – per unit mass the weaker-looking alloy beats the steel, and other criteria such as cost and corrosion can outweigh strength entirely.

Key:

From models to materials: bonding models predict bulk properties – metallic conducts and bends, ionic is hard and brittle, giant covalent resists heat, molecular and polymer materials melt low and flex. Material selection compares measured properties against a stated requirement, often as a ratio such as strength per density.

Structure 3.1 · The periodic table: classification of elements

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout covers selected HL/SL topics for the course first assessed in 2025; it is not the whole syllabus.

Prerequisites: electron configurations from Structure 1.3 and the nuclear-atom model from Structure 1.2.

The examples are original and fictional: they are not real measurements or real student investigations.

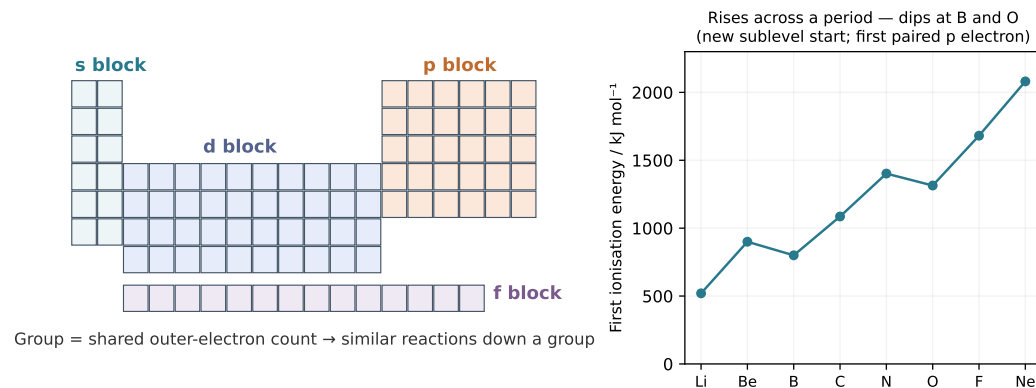
The periodic table: classification of elements

■ What would explain this observation?

- Lithium, sodium and potassium all fizz with water – and each more violently than the last. Why do these elements behave like a family?

■ Build the model

- Elements are ordered by atomic number. A vertical **group** is a column of elements. Main-group elements usually share an outer-electron count within a group and show related chemistry. A horizontal **period** is a row of the table. The **block** (s, p, d or f) names which sublevel is being filled.
- Across a period the **first ionisation energy** rises overall, because the nuclear charge grows at almost unchanged shielding; it dips where a new sublevel starts (B after Be) and where a p orbital first holds a paired electron (O after N). Down a group it falls as the outer electron is farther away and more shielded.
- group: A vertical column of the periodic table; period: A horizontal row of the periodic table; block: The region of the table named by the sublevel being filled; first ionisation energy: The energy needed to remove one electron from each atom in one mole of gaseous atoms.



Ionization energy rises across a period, with dips at B (new p sublevel) and O (first paired p electron)

■ Choose evidence that can test it

- Compare the reactions of two or more members of one group under the same conditions, such as lithium and potassium with cold water: same products, different vigour – evidence that shared outer configuration drives the chemistry.
- Plot measured first ionisation energies across a period. The model predicts the rise and the two dips at the sublevel boundary and the first paired p electron; check whether the data show exactly those dips.

■ Work from known quantities

- Known: the first ionisation energies of magnesium and aluminium are 738 and 578 kJ/mol. The trend rises across a period, yet Al lies below Mg. The electron removed from aluminium is in 3p, while magnesium loses from 3s: the 3p electron is higher in energy and better shielded, so less energy is needed – a difference of $738 - 578 = 160$ kJ/mol.

Example:

The first ionisation energies of sodium and magnesium are 496 and 738 kJ/mol. Find the difference and explain its direction using the same sequence: known quantities → model → relation → substitution → unit and interpretation. (Answer: $738 - 496 = 242$ kJ/mol higher for magnesium; the 3s electron in Mg feels one more proton at nearly the same shielding, so it is held more tightly.)

■ Check the conclusion and its limits

- The block model explains the MAIN trends; transition-metal chemistry has additional d-orbital detail beyond this classification lesson, and ionisation values are measured for gaseous atoms, not for reacting solids.

Warn:

Elements in the same group react alike because they have the same number of electrons. This claim is false: they share the number of OUTER electrons – sodium has 11 electrons in total and potassium 19, yet both react as Group 1 metals because each has one outer electron.

Key:

The periodic table: classification of elements: group = shared outer-electron count and similar reactions; period = a filling shell; block = the sublevel being filled. Ionisation energy rises across a period with dips at the sublevel boundary and the first paired p electron, and falls down a group.

Functional groups: classification of organic compounds

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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Prerequisites: read the stated quantities and units, and use arithmetic and the model conditions below.

The examples are original and fictional: they are not real measurements or real student investigations.

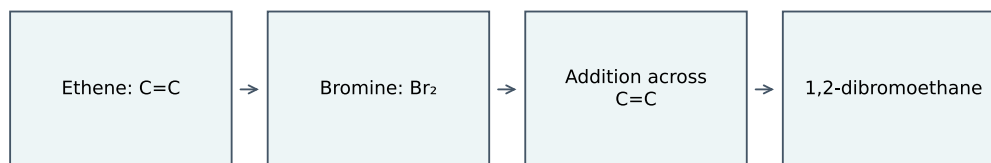
Organic structures and reaction pathways

■ What would explain this observation?

- Two compounds can have the same molecular formula but different structures. Their **functional groups** help predict which reactions they undergo.

■ Build the model

- A homologous series shares a functional group and general formula.
- Structural **isomers** share a molecular formula but differ in atom connections.
- Alkenes contain a carbon-carbon double bond.
- functional group: An atom group determining characteristic reactions.
- isomer : A compound sharing a formula but differing in structure.



Organic structures and reaction pathways: original worked-case diagram

■ Choose evidence that can test it

- Distinguish addition, substitution, oxidation and polymerization by tracing bonds before and after reaction.
- Conditions and reagents belong to the reaction arrow; they are not interchangeable labels.

In the lab: Draw displayed or structural formulae with the correct number of bonds at each carbon. Use a carbon count to check a proposed synthesis. At advanced level, track reagents and conditions through multistep routes.

■ Work from known quantities

- Known: ethene adds bromine across its double bond.
- The two-carbon skeleton stays intact and each carbon gains one bromine atom, giving 1,2-dibromoethane.
- One mole of ethene reacts with one mole of bromine in this addition reaction.

Your turn: How many moles of bromine react completely with 0.15 mol ethene?

■ Check the conclusion and its limits

- Bromine decolourization provides evidence of unsaturation in an appropriate test.
- It is not proof that an unknown sample is specifically ethene.

Not true: Compounds with the same molecular formula always have identical properties. Same molecular formula is not same compound: isomers such as butane and 2-methylpropane share C_4H_{10} but have different structures and different physical properties. Structure, not formula alone, sets properties.

Key:

- A homologous series shares a functional group and general formula.
- Distinguish addition, substitution, oxidation and polymerization by tracing bonds before and after reaction.
- Watch for: bromine decolourization provides evidence of unsaturation in an appropriate test.

■ Write from 0.25 mol of bromine

- Ethene adds bromine in a 1 to 1 mole ratio.
- Find the bromine amount that reacts with 0.25 mol of ethene. The same molecular formula can belong to different compounds.

Measuring enthalpy change

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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Prerequisites: read the stated quantities and units, and use arithmetic and the model conditions below.

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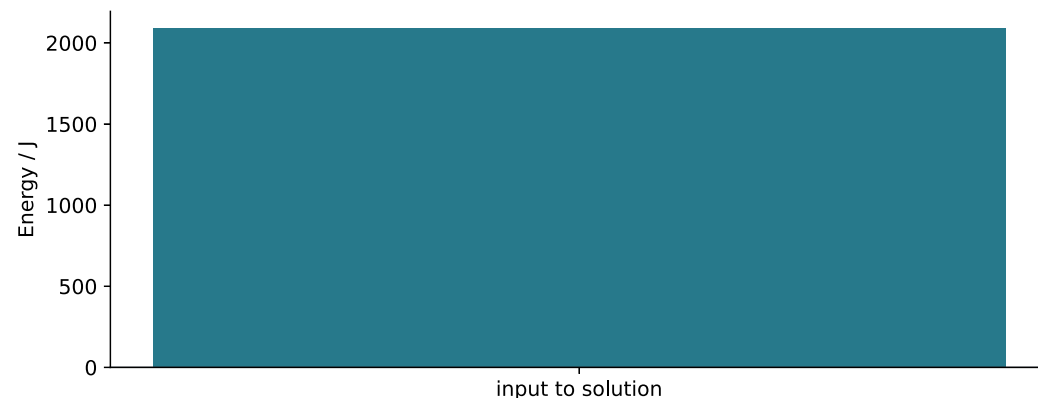
Calorimetry and chemical energy

■ What would explain this observation?

- A cup warms when two solutions react. The temperature rise measures energy transferred to the surroundings; it does not directly equal the **enthalpy change**.

■ Build the model

- **Exothermic** reactions transfer energy to surroundings.
- Endothermic reactions take energy from surroundings.
- Bond breaking requires energy; bond formation releases energy.
- exothermic: Transferring energy to the surroundings.
- enthalpy change: Heat change at constant pressure for a stated process.



Calorimetry and chemical energy: original worked-case diagram

■ Choose evidence that can test it

- Use energy transferred = mass $\times$ specific heat capacity $\times$ temperature change.
- Convert joules to kilojoules before dividing by reaction amount.
- An exothermic molar enthalpy change has a negative sign.

In the lab: Use insulation and a lid, measure starting temperatures consistently, stir, and record a temperature-time series. Estimate the reaction temperature from an appropriate extrapolation rather than ignoring cooling during measurement.

■ Work from known quantities

- Known: 100 g solution rises by 5.0 °C; specific heat capacity is 4.18 J per gram per degree. $q = mc\Delta T$. $q = 100 \times 4.18 \times 5.0 = 2,090 \text{ J} = 2.09 \text{ kJ}$.
- If 0.050 mol reacts, $\Delta H = -q/n = -2.09/0.050 = -41.8 \text{ kJ per mole}$.

Your turn: 50 g water rises 4 °C. Use $c = 4.2 \text{ J per gram per degree}$ to find q .

■ Check the conclusion and its limits

- Heat loss usually lowers the observed temperature rise.
- The solution gaining heat and the reaction losing heat have opposite signs.

Not true: Bond breaking releases energy. Bond breaking is endothermic - it requires energy. Bond forming releases energy. Confusing the two reverses the sign of every energy calculation.

Key:

- Exothermic reactions transfer energy to surroundings.
- Use energy transferred = mass $\times$ specific heat capacity $\times$ temperature change.
- Watch for: heat loss usually lowers the observed temperature rise.

■ Write from 756 J and minus 37.8 kJ per mole

- 60 g of solution rises by 3 degrees. Use $c = 4.2 \text{ J per gram per degree}$. 0.020 mol reacts.
- Find q and the enthalpy change per mole. Bond breaking requires energy.

Reactivity 1.2 · Energy cycles in reactions

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout covers selected HL/SL topics for the course first assessed in 2025; it is not the whole syllabus.

Prerequisites: enthalpy change and calorimetry from Reactivity 1.1, balancing equations, and arithmetic with signed quantities.

The examples are original and fictional: they are not real measurements or real student investigations.

Energy cycles in reactions

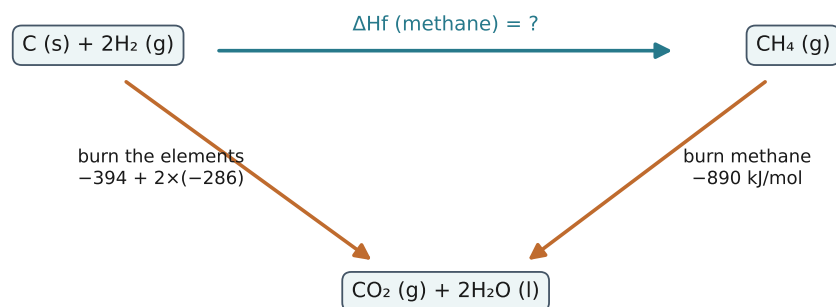
■ What would explain this observation?

- Some enthalpy changes cannot be measured directly: carbon and hydrogen will not simply combine into methane in a calorimeter. How can we still know the enthalpy change of that reaction?
-

■ Build the model

- **Hess's law** : the total enthalpy change of a reaction is the same whatever route is taken, because energy is conserved. An **energy cycle** links the direct route to an indirect route through shared substances, so the unknown change is found from measurable ones.
- With **standard enthalpy of formation** data the rule is $\Delta H = \Sigma\Delta H_f(\text{products}) - \Sigma\Delta H_f(\text{reactants})$. With **standard enthalpy of combustion** data the cycle runs through the combustion products: $\Delta H = \Sigma\Delta H_c(\text{reactants}) - \Sigma\Delta H_c(\text{products})$. Elements in their standard states have $\Delta H_f = 0$.
- energy cycle: A diagram linking two routes between the same start and end points; standard enthalpy of formation: The enthalpy change when one mole of a compound forms from its elements under standard conditions; standard enthalpy of combustion: The enthalpy change when one mole burns completely under standard conditions.

Hess cycle: same start, same end → same total energy change



Energy cycles in reactions: original worked-case diagram

■ Choose evidence that can test it

- Measure the enthalpies of combustion of carbon, hydrogen and methane by calorimetry; each is directly measurable. The cycle then predicts the formation enthalpy of methane, and the same method applied to a reaction that CAN be measured directly (such as neutralisation) confirms that route and direct measurement agree within experimental uncertainty.
- State the states of every substance and the standard conditions; an enthalpy value without states is incomplete.

■ Work from known quantities

- Known: $\Delta H_c(\text{C}) = -394 \text{ kJ/mol}$, $\Delta H_c(\text{H}_2) = -286 \text{ kJ/mol}$, $\Delta H_c(\text{CH}_4) = -890 \text{ kJ/mol}$. The target is ΔH_f for $\text{C (s)} + 2\text{H}_2 \text{ (g)} \rightarrow \text{CH}_4 \text{ (g)}$. By Hess's law along the combustion route: $\Delta H_f = [\Delta H_c(\text{C}) + 2\Delta H_c(\text{H}_2)] - \Delta H_c(\text{CH}_4)$. Substitution: $\Delta H_f = [-394 + 2 \times (-286)] - (-890) = -966 + 890 = -76 \text{ kJ/mol}$. The negative sign means the formation of methane from its elements is exothermic.

Example:

Ethanol burns: $\text{C}_2\text{H}_5\text{OH (l)} + 3\text{O}_2 \text{ (g)} \rightarrow 2\text{CO}_2 \text{ (g)} + 3\text{H}_2\text{O (l)}$. Using $\Delta H_f(\text{CO}_2) = -394$, $\Delta H_f(\text{H}_2\text{O}) = -286$ and $\Delta H_f(\text{C}_2\text{H}_5\text{OH}) = -278 \text{ kJ/mol}$, find ΔH_c of ethanol with the same sequence: known quantities → model → relation → substitution → unit and interpretation. (Answer: $\Delta H = [2 \times (-394) + 3 \times (-286)] - [-278 + 0] = -1646 + 278 = -1368 \text{ kJ/mol}$; oxygen is an element in its standard state, so its ΔH_f is 0.)

■ Check the conclusion and its limits

- Hess's law assumes the routes share the same start and end states and that the tabulated values apply at the stated conditions. Real calorimetry loses heat, so measured values carry uncertainty; the law itself is exact.

Warn:

The enthalpy change of a reaction depends on the route by which the reaction happens. This claim is false: Hess's law follows from the conservation of energy – the total change depends only on the start and end states, so an unmeasurable change can be computed from a measurable indirect route.

Key:

Energy cycles in reactions: ΔH is route-independent. Use $\Delta H = \Sigma \Delta H_f(\text{products}) - \Sigma \Delta H_f(\text{reactants})$ from formation data, or $\Delta H = \Sigma \Delta H_c(\text{reactants}) - \Sigma \Delta H_c(\text{products})$ from combustion data, keeping signs and states explicit.

Energy from fuels

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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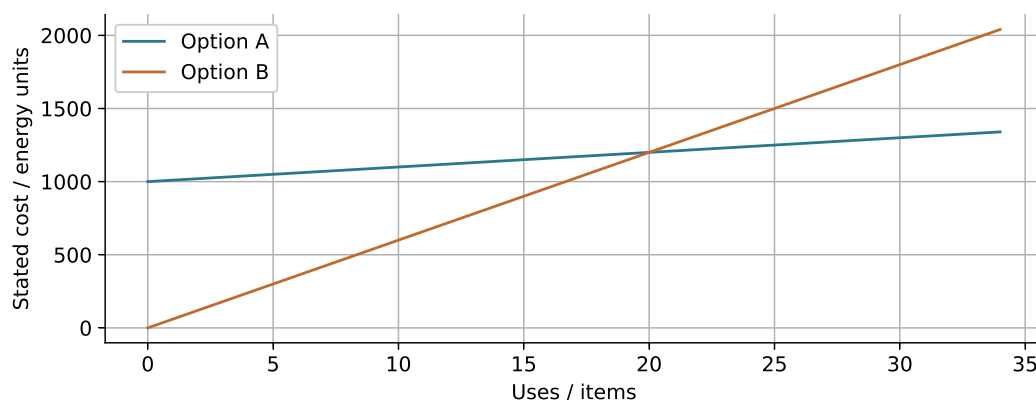
Atmosphere, resources and life-cycle decisions

■ What would explain this observation?

- A reusable container can require more energy to manufacture than a single-use one. Its impact depends on how often it is used and what happens at disposal.

■ Build the model

- A **life-cycle assessment** considers raw materials, manufacture, transport, use and end-of-life processes.
- Greenhouse gases absorb and emit infrared radiation; pollution and climate effects are related but distinct questions.
- life-cycle assessment: Assessment across production, use and disposal.
- **functional unit** : The common service used for a fair comparison.



Atmosphere, resources and life-cycle decisions: original worked-case diagram

■ Choose evidence that can test it

- Define the functional unit before comparing products.
- The same delivered service, such as carrying one litre of water a hundred times, is fairer than comparing one object with another regardless of lifetime.

In the lab: List system boundaries, energy sources and assumptions. Compare water demand, emissions and waste separately before making a judgement. Explain whose priorities affect the decision and where the data are uncertain.

■ Work from known quantities

- Known: a reusable item requires 1,000 energy units to make and 10 per use; a disposable item requires 60 per use.
- Equality occurs when $1,000 + 10n = 60n$.
- Rearranging gives $1,000 = 50n$, so $n = 20$ uses under this simplified model.

Your turn: A reusable item costs 600 energy units initially and 5 per use; disposables cost 35 per use. Find break-even uses.

■ Check the conclusion and its limits

- A break-even model is sensitive to cleaning, transport and disposal assumptions.
- Recyclable does not guarantee that an item is actually collected and recycled.

Not true: A recyclable product is always recycled after use. Recyclable means the material CAN be recycled with the right collection and processing; actual recycling depends on local systems and behaviour. Many recyclable items are never collected.

Key:

- A life-cycle assessment considers raw materials, manufacture, transport, use and end-of-life processes.
- Define the functional unit before comparing products.
- Watch for: A break-even model is sensitive to cleaning, transport and disposal assumptions.

■ Write from a break-even of 30 uses

- A reusable item costs 900 energy units to make and 6 per use. A disposable item costs 36 per use.
- Find the number of uses where the totals are equal. Recyclable does not mean the item is recycled.

Entropy and spontaneity

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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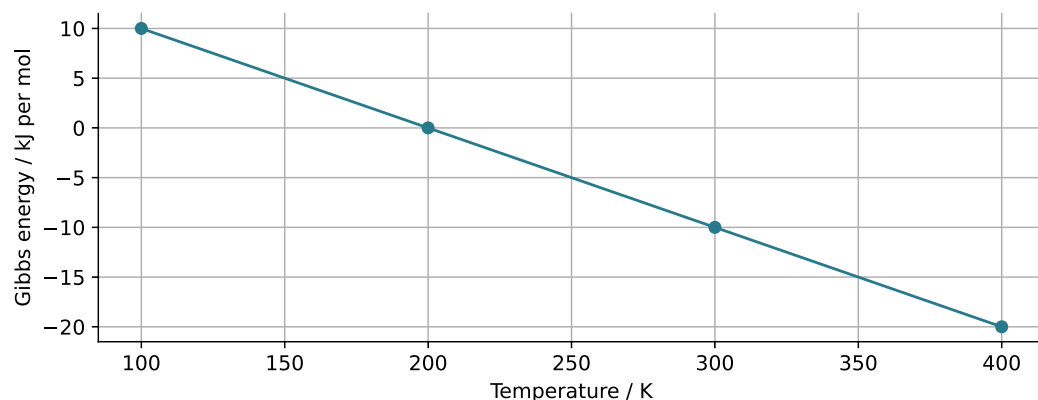
Entropy, Gibbs energy and feasibility

■ What would explain this observation?

- A reaction can be endothermic and still be thermodynamically favourable. Enthalpy alone does not determine the direction favoured at a given temperature.

■ Build the model

- **Entropy** is associated with energy dispersal and the number of accessible microscopic arrangements.
- **Gibbs energy** combines enthalpy, entropy and absolute temperature for a process at constant temperature and pressure.
- entropy: A state property related to energy dispersal and accessible arrangements.
- Gibbs energy: A thermodynamic quantity combining enthalpy and entropy contributions.



Entropy, Gibbs energy and feasibility: original worked-case diagram

■ Choose evidence that can test it

- Use $\Delta G = \Delta H - T\Delta S$ with consistent energy units.
- A negative Gibbs energy change indicates thermodynamic favourability for the stated conditions, not a fast rate.
- An activation barrier can make a favourable process slow.

In the lab: State whether values are standard-state quantities and record temperature in kelvin. Convert entropy from joules per kelvin per mole into kilojoules per kelvin per mole when enthalpy is in kilojoules per mole.

■ Work from known quantities

- Known: $\Delta H = +20$ kJ per mole, $\Delta S = +100$ J per kelvin per mole and $T = 300$ K.
- Convert $\Delta S = 0.100$ kJ per kelvin per mole. $\Delta G = \Delta H - T\Delta S = 20 - 300 \times 0.100 = -10$ kJ per mole.
- The process is favourable under the stated approximation.

Your turn: $\Delta H = 30$ kJ per mole, $\Delta S = 0.10$ kJ per kelvin per mole and $T = 400$ K. Find ΔG .

■ Check the conclusion and its limits

- Thermodynamic favourability does not establish reaction rate.
- Celsius cannot replace kelvin in $T\Delta S$, and a unit conversion error can change the result by a factor of 1,000.

Not true: A thermodynamically favourable reaction must be fast. Feasibility and speed are different questions. Thermodynamics (ΔG) says a reaction can proceed; kinetics (activation energy, rate) says how fast. Diamond to graphite is favourable yet imperceptibly slow.

Key:

- Entropy is associated with energy dispersal and the number of accessible microscopic arrangements.
- Use $\Delta G = \Delta H - T\Delta S$ with consistent energy units.
- Watch for: thermodynamic favourability does not establish reaction rate.

■ Write from a Gibbs energy change of plus 8 kJ per mole

- Delta H is plus 40 kJ per mole. Delta S is plus 80 J per kelvin per mole. The temperature is 400 K.
- Find delta G. A favourable reaction is not necessarily fast.

How much? The amount of chemical change

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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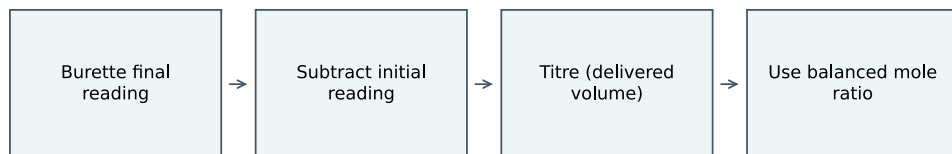
Titration and a defensible concentration

■ What would explain this observation?

- A burette reading is not the delivered volume. The **titre** is the difference between final and initial readings, and both readings have uncertainty.

■ Build the model

- A titration measures the amount of one solution needed to react with a known amount of another.
- The equation gives the mole ratio.
- An indicator endpoint approximates the **equivalence point** when a suitable indicator is used.
- titre: The volume delivered between two burette readings.
- equivalence point: The point of stoichiometric reaction completion.



■ Choose evidence that can test it

- Calculate the known amount first, apply the stoichiometric ratio, then divide by the unknown solution volume in cubic decimetres.
- Use concordant titres as required by the school method and report the accepted values.

In the lab: Rinse the burette with its solution and the pipette with the solution it transfers. Rinse the flask with distilled water. Add titrant slowly near the end-point, swirl, and read the meniscus at eye level. Use a white tile and appropriate eye protection.

■ Work from known quantities

- Known: 25.0 cubic centimetres of acid reacts 1:1 with 20.0 cubic centimetres of 0.100 mol per cubic decimetre alkali. $n = cV = 0.100 \times 0.0200 = 0.00200$ mol.
- Acid amount is 0.00200 mol. $c = n/V = 0.00200/0.0250 = 0.0800$ mol per cubic decimetre.

Your turn: Burette readings are 1.40 and 23.65 cubic centimetres. Calculate titre.

■ Check the conclusion and its limits

- Adding distilled water to the flask changes volume but not the transferred amount of analyte.
- Do not average a rough titre with carefully measured concordant values.

Not true: A flask must be rinsed with analyte to avoid dilution of its mole amount. The pipette delivers a fixed AMOUNT of analyte into the flask; extra water in the flask changes the volume but not the moles present, and the titration measures moles. Rinsing the flask with analyte would add unknown extra moles and spoil the result.

Key:

- A titration measures the amount of one solution needed to react with a known amount of another.
- Calculate the known amount first, apply the stoichiometric ratio, then divide by the unknown solution volume in cubic decimetres.
- Watch for: adding distilled water to the flask changes volume but not the transferred amount of analyte.

■ Write from a titre of 20.50 cubic centimetres

- Burette readings are 0.85 and 21.35 cubic centimetres.
- Find the titre. Extra water in the flask does not change the amount delivered by the pipette.

■ Write from a mean titre

- Three concordant titres: 20.50, 20.60 and 20.40 cubic centimetres.
- Mean = 20.50 cubic centimetres – only CONCORDANT runs (within 0.10) enter the mean; a wild 23.10 is discarded with a reason.

■ Write from a one-in-five dilution

- 25.0 cubic centimetres of stock is diluted to 250.0 in a volumetric flask; 25.0 of THAT is diluted to 250.0 again.
- Each step dilutes tenfold, so the final concentration is $\text{stock} \div 100$ – a dilution factor multiplies along the chain.

How fast? The rate of chemical change

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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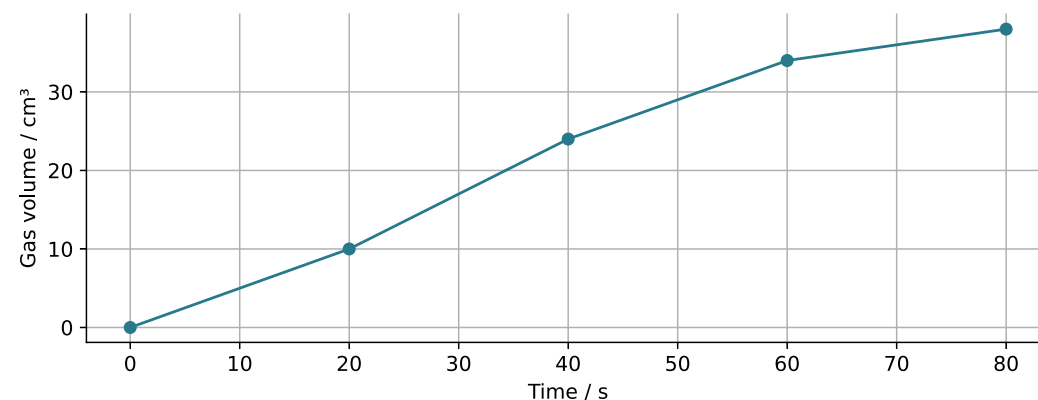
Rates, catalysts and reliable endpoints

■ What would explain this observation?

- A faster reaction finishes sooner, but it need not make more product. **Rate** and final yield answer different questions.

■ Build the model

- Reaction rate describes reactant used or product formed per time.
- Higher temperature increases the fraction of collisions with enough energy.
- A catalyst provides an alternative pathway with lower **activation energy**.
- activation energy: The energy barrier for a reaction pathway.
- rate: Change in a measured quantity per unit time.



Rates, catalysts and reliable endpoints: original worked-case diagram

■ Choose evidence that can test it

- A product-time graph has a steeper gradient where rate is larger.
- A tangent estimates instantaneous rate; a secant gives average rate over an interval.
- The final plateau reflects the total collected product under the stated conditions.

In the lab: For gas production, check apparatus for leaks, start timing consistently and record volume at regular intervals. Keep concentration, reactant amount and surface area controlled when changing temperature.

■ Work from known quantities

- Known: gas volume increases from 10 to 34 cubic centimetres between 20 and 60 s.
- Average rate = change in volume/change in time.
- Rate = $(34 - 10)/(60 - 20) = 0.60$ cubic centimetres per second.
- This is not necessarily the instantaneous rate at 40 s.

Your turn: A reaction makes 24 cubic centimetres of gas in 40 s. Find average rate.

■ Check the conclusion and its limits

- A catalyst does not change the equilibrium constant at a fixed temperature.
- A mass-loss method cannot detect all reactions, and losing gas through a leak biases a collection experiment.

Not true: A faster reaction always produces a larger final amount of product. Rate measures how quickly product forms, not how much finally forms. The final amount is fixed by the limiting reagent; a catalyst or powder changes only the time taken to reach it.

Key:

- Reaction rate describes reactant used or product formed per time.
- A product-time graph has a steeper gradient where rate is larger.
- Watch for: A catalyst does not change the equilibrium constant at a fixed temperature.

■ Write from an average rate of 0.40 cubic centimetres per second

- Gas volume rises from 8 to 20 cubic centimetres between 10 s and 40 s.
- Find the average rate. A faster rate does not increase the final amount.

How far? The extent of chemical change

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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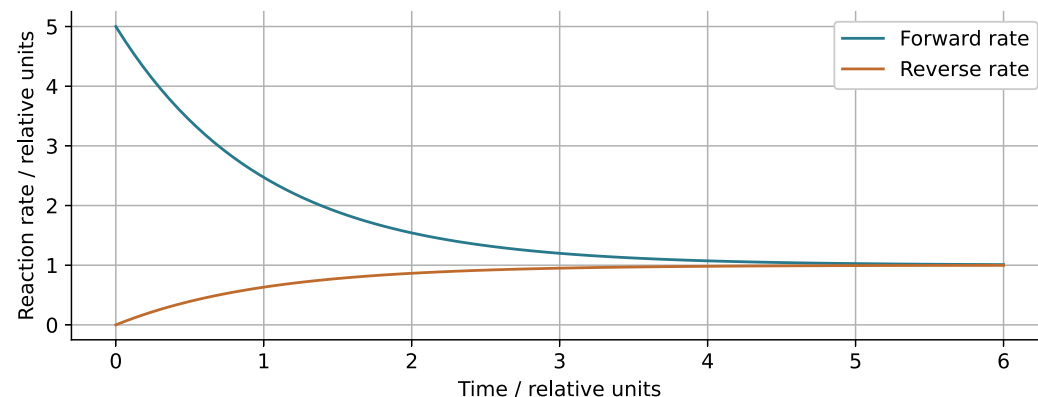
Equilibrium and changing conditions

■ What would explain this observation?

- A **reversible reaction** can continue in a closed vessel while measured concentrations stay constant. Constant composition does not mean particles have stopped reacting.

■ Build the model

- Dynamic **equilibrium** occurs in a closed system when forward and reverse rates are equal.
- Reactant and product concentrations are constant, but they need not be equal.
- equilibrium: A state with equal forward and reverse reaction rates.
- reversible reaction: A reaction that can proceed in both directions.



Equilibrium and changing conditions: original worked-case diagram

■ Choose evidence that can test it

- A concentration or pressure change disturbs the balance.
- The system responds toward a new equilibrium.
- Temperature changes can also change the equilibrium constant; a catalyst changes how quickly equilibrium is reached.

In the lab: State the balanced equation and whether the forward reaction is exothermic before predicting a temperature effect. Count gas coefficients when considering pressure; pressure has no composition effect when gaseous amounts are equal on both sides.

■ Work from known quantities

- Known: in $A \rightleftharpoons B$, equilibrium concentrations are $[A] = 0.20$ and $[B] = 0.60$ in the same concentration unit.
- For this stated expression, $K = [B]/[A]$.
- $K = 0.60/0.20 = 3.0$.
- Equal rates do not imply $K = 1$.

Your turn: For $A \rightleftharpoons B$, calculate $[B]/[A]$ when $[B]=0.8$ and $[A]=0.2$.

■ Check the conclusion and its limits

- Do not use a catalyst to claim a larger equilibrium yield.
- For heterogeneous equilibria, pure solids are omitted from the usual equilibrium expression.

Not true: A catalyst changes the equilibrium constant at a fixed temperature. A catalyst speeds the forward and reverse reactions equally, so equilibrium is reached sooner but the equilibrium constant and final composition at a fixed temperature are unchanged.

Key:

- Dynamic equilibrium occurs in a closed system when forward and reverse rates are equal.
- A concentration or pressure change disturbs the balance.
- Watch for: do not use a catalyst to claim a larger equilibrium yield.

■ Write from an equilibrium constant of 2.5

- For A in equilibrium with B, the concentration of B is 1.00 and the concentration of A is 0.40, in the same unit.
- Find K for the stated expression of B divided by A. A catalyst does not change K at a fixed temperature.

Reactivity 3.1 · Proton transfer reactions

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout covers selected HL/SL topics for the course first assessed in 2025; it is not the whole syllabus.

Prerequisites: chemical equations and reversible reactions from Structure and Reactivity 2, logarithms, and the mole.

The examples are original and fictional: they are not real measurements or real student investigations.

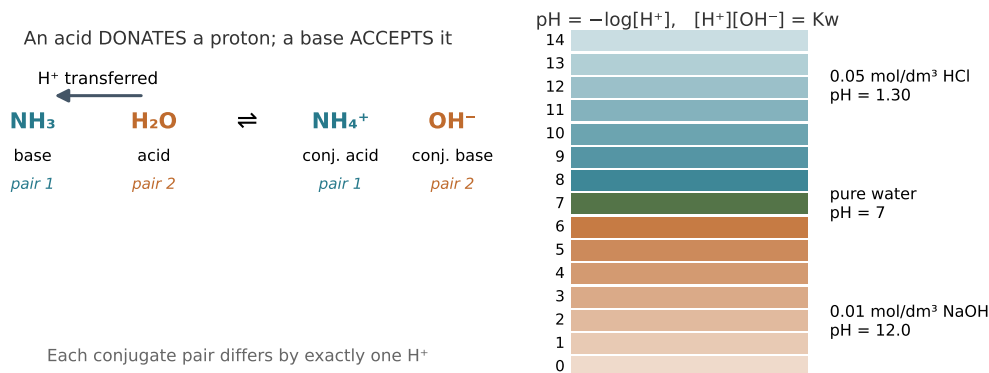
Proton transfer reactions

■ What would explain this observation?

- Two bottles both labelled “0.1 mol/dm³ acid” give very different pH readings: 1.0 and 2.9. Same concentration, different acidity – how?
-

■ Build the model

- In the Brønsted–Lowry model, an acid is a **proton donor** and a base is a proton acceptor. Every such reaction links a **conjugate acid–base pair** differing by exactly one H⁺: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ pairs NH₃ with NH₄⁺ and H₂O with OH[−].
- A **strong acid** dissociates almost completely in water; a **weak acid** only partially, so the same concentration gives a different [H⁺]. Water itself ionises slightly: $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ mol}^2/\text{dm}^6$ at 298 K, and $\text{pH} = -\log[\text{H}^+]$.
- proton donor: A species that gives an H⁺ to another species; conjugate acid–base pair: Two species differing by exactly one proton; strong acid: An acid that dissociates almost completely in water; weak acid: An acid that only partially dissociates in water.



Proton transfer reactions: original worked-case diagram

■ Choose evidence that can test it

- Compare equal-concentration solutions: a strong acid conducts better and shows a lower pH than a weak acid of the same concentration, because more of its molecules have dissociated. Measure pH with a calibrated meter, not by estimating from the label.
- Adding the conjugate base of a weak acid shifts its equilibrium toward the undissociated form – evidence that the dissociation is a genuine equilibrium, not a one-way reaction.

■ Work from known quantities

- Known: HCl is a strong acid at 0.05 mol/dm³, so $[\text{H}^+] = 0.05 \text{ mol/dm}^3$. $\text{pH} = -\log[\text{H}^+]$. Substitution: $\text{pH} = -\log(0.05) = 1.30$. For 0.01 mol/dm³ NaOH, a strong base, $[\text{OH}^-] = 0.01$; using K_w , $[\text{H}^+] = K_w/[\text{OH}^-] = 1.0 \times 10^{-14} / 0.01 = 1.0 \times 10^{-12}$, so $\text{pH} = 12.0$.

Example:

A strong-acid solution has $[\text{H}^+] = 2.5 \times 10^{-4} \text{ mol/dm}^3$. Find its pH using the same sequence: known quantities → model → relation → substitution → unit and interpretation. (Answer: $\text{pH} = -\log(2.5 \times 10^{-4}) = 3.60$. pH is dimensionless; state the input concentration with its unit.)

■ Check the conclusion and its limits

- $\text{pH} = -\log[\text{H}^+]$ in this form assumes dilute solutions where concentration approximates activity; at high concentration the approximation weakens. $K_w = 1.0 \times 10^{-14}$ holds at 298 K and changes with temperature.

Warn:

A strong acid is the same thing as a concentrated acid. This claim is false: strength is the degree of dissociation and concentration is the amount per volume; a concentrated weak acid and a dilute strong acid both exist.

Key:

Proton transfer reactions: acids donate protons and bases accept them; every transfer links a conjugate pair differing by one H⁺. Strength (degree of dissociation) is not concentration, and $\text{pH} = -\log[\text{H}^+]$ with $K_w = [\text{H}^+][\text{OH}^-]$.

■ Write from pH 3

- A solution has hydrogen-ion concentration $1 \times 10^{-3} \text{ mol dm}^{-3}$.
- Find the pH: $-\log(10^{-3}) = 3$.

■ Write from pH 11

- A solution has pH 11.
- Find both ion concentrations: $[\text{H}^+] = 10^{-11}$ and $[\text{OH}^-] = K_w \div [\text{H}^+] = 10^{-3}$ at 25 °C.

■ Write from the strength trap

- Solution A is a concentrated weak acid; solution B is a dilute strong acid.
- Which can have the lower pH? Either – strength is dissociation, concentration is amount per volume.

Electron transfer reactions

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout covers selected HL topics for the course first assessed in 2025; it is not the whole syllabus.

Prerequisites: read the stated quantities and units, and use arithmetic and the model conditions below.

The examples are original and fictional: they are not real measurements or real student investigations.

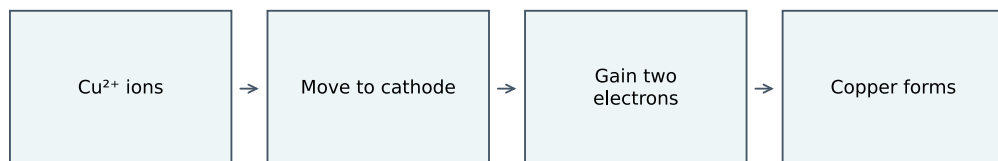
Redox and electrolysis

■ What would explain this observation?

- An aqueous salt solution can produce different electrode products from the molten salt. Water introduces competing species into the system.

■ Build the model

- **Oxidation** is loss of electrons and **reduction** is gain of electrons.
- In electrolysis, cations move toward the cathode and anions toward the anode.
- Reduction occurs at the cathode.
- oxidation: Loss of electrons.
- reduction: Gain of electrons.



Redox and electrolysis: original worked-case diagram

■ Choose evidence that can test it

- Predict products using the specified electrolyte and electrode material.
- In an aqueous solution, hydrogen or oxygen may form because water-related species compete.
- Molten salts contain only the ions of the salt.

In the lab: Use a low-voltage direct-current supply, approved electrodes, and the school risk assessment. Collect gases only by an approved method. Keep chlorine demonstrations teacher-controlled; do not ask students to generate hazardous gases independently.

■ Work from known quantities

- Known: a copper ion gains two electrons.
- Half-equation: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$.
- One mole of Cu^{2+} requires two moles of electrons.
- For 0.050 mol of copper, electron amount = $2 \times 0.050 = 0.100$ mol.

Your turn: How many moles of electrons reduce 0.20 mol of Cu^{2+} to copper?

■ Check the conclusion and its limits

- Electrode signs depend on the cell type.
- In an electrolytic cell the cathode is negative; reduction remains the defining process at a cathode in every cell.

Not true: Reduction means loss of electrons. Reduction is defined as gain of electrons (OILRIG: oxidation is loss, reduction is gain). At a cathode, reduction occurs in every cell type, whatever the electrode sign.

Key:

- Oxidation is loss of electrons and reduction is gain of electrons.
- Predict products using the specified electrolyte and electrode material.
- Watch for: electrode signs depend on the cell type.

■ Write from 0.50 mol of electrons

- The half-equation is $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$.
- Find the electron amount that reduces 0.25 mol of copper(II) ions. Reduction is gain of electrons.

Reactivity 3.3 · Electron sharing reactions

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout covers selected HL/SL topics for the course first assessed in 2025; it is not the whole syllabus.

Prerequisites: covalent bonding from Structure 2.2 and alkane structure from Structure 3.2.

The examples are original and fictional: they are not real measurements or real student investigations.

Electron sharing reactions

■ What would explain this observation?

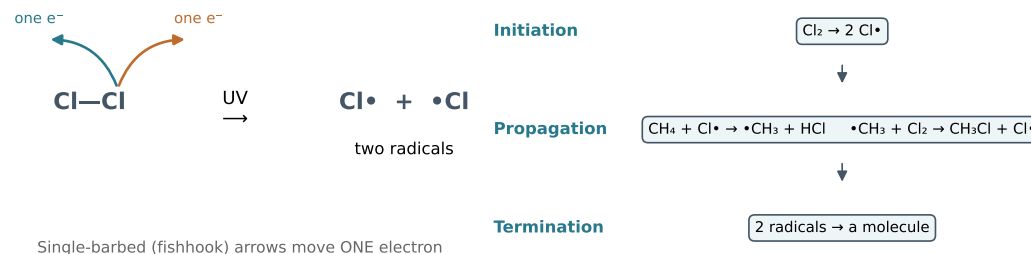
- Methane and chlorine are mixed in the dark and nothing happens; the same mixture in sunlight reacts at once, and the product is never pure chloromethane – it always carries further-substituted molecules. Why?

■ Build the model

- A covalent bond can break by **homolytic fission** : the shared pair splits so each atom keeps one electron, forming two **free radicals** – atoms or groups with an unpaired electron, written with a dot, such as $\text{Cl}\cdot$. A single-barbed (fishhook) arrow shows ONE electron moving.
- Radical substitution runs as a chain: **initiation** (UV light splits Cl_2 into two $\text{Cl}\cdot$), **propagation** (a radical makes a product and another radical: $\text{CH}_4 + \text{Cl}\cdot \rightarrow \cdot\text{CH}_3 + \text{HCl}$, then $\cdot\text{CH}_3 + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \text{Cl}\cdot$), and **termination** (two radicals join, ending the chain).
- homolytic fission: Bond breaking that gives each atom one electron of the shared pair; free radical: A species with an unpaired electron, highly reactive; initiation: The step that first produces radicals; propagation: A chain step consuming one radical and producing another; termination: A step removing radicals by joining two of them.

Homolytic fission: one electron to each atom

Radical chain substitution



Electron sharing reactions: original worked-case diagram

■ Choose evidence that can test it

- The reaction needs UV light or heat to start but then continues on its own – evidence for a chain in which each step regenerates a radical. Trace amounts of ethane among the products show that two $\cdot\text{CH}_3$ radicals joined in a termination step.
- A mixture of CH_3Cl , CH_2Cl_2 , CHCl_3 and CCl_4 forms because the product still has C–H bonds open to further propagation; single-substitution purity is evidence against a controlled, non-chain route.

■ Work from known quantities

- Known: the bond enthalpies are $\text{Cl}-\text{Cl}$ 242 kJ/mol and $\text{C}-\text{H}$ 413 kJ/mol. Bond breaking is endothermic, so the weaker bond breaks first. Comparing the two: $413 - 242 = 171$ kJ/mol less energy breaks $\text{Cl}-\text{Cl}$, which is why UV initiates the chain by splitting chlorine rather than methane.

Example:

The $\text{H}-\text{Cl}$ bond enthalpy is 431 kJ/mol and the $\text{C}-\text{Cl}$ bond in CH_3Cl is 339 kJ/mol. For the propagation step $\text{CH}_4 + \text{Cl}\cdot \rightarrow \cdot\text{CH}_3 + \text{HCl}$, estimate the energy change using the same sequence: known quantities $\rightarrow$ model $\rightarrow$ relation $\rightarrow$ substitution $\rightarrow$ unit and interpretation. (Answer: break $\text{C}-\text{H}$ (+413), form $\text{H}-\text{Cl}$ (−431): $\Delta\text{H} \approx 413 - 431 = -18$ kJ/mol – slightly exothermic, so the step proceeds readily once radicals exist.)

■ Check the conclusion and its limits

- Bond enthalpies are averages measured over many compounds, so an estimate from them is approximate, not exact. The chain model for this reaction explains the UV start and the product mixture; it does not describe polar reagents, where heterolytic (electron-PAIR) mechanisms apply instead – the next lesson’s model.

Warn:

In a radical substitution, every chlorine atom ends up substituted onto a methane molecule. This claim is false: termination joins radicals with each other ($\text{Cl}\cdot + \text{Cl}\cdot \rightarrow \text{Cl}_2$, or two $\cdot\text{CH}_3 \rightarrow \text{C}_2\text{H}_6$), and propagation continues substituting the products, so the outcome is always a mixture.

Key:

Electron sharing reactions: homolytic fission splits a shared pair one electron each way, giving radicals; fishhook arrows track single electrons. Radical substitution chains through initiation, propagation and termination, which explains the UV trigger and the mixture of products.

Reactivity 3.4 · Electron-pair sharing reactions

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout covers selected HL/SL topics for the course first assessed in 2025; it is not the whole syllabus.

Prerequisites: covalent bonding and polarity from Structure 2.2, alkenes and halogenoalkanes from Structure 3.2, and radical (electron-sharing) mechanisms from Reactivity 3.3.

The examples are original and fictional: they are not real measurements or real student investigations.

Electron-pair sharing reactions

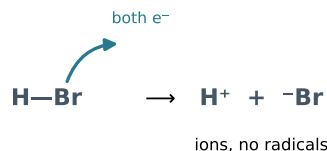
■ What would explain this observation?

- Bromoethane shaken with warm sodium hydroxide solution turns into ethanol, while the same alkane skeleton with no C–Br bond ignores the hydroxide completely. What does the C–Br bond make possible?

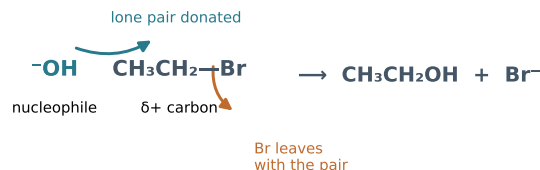
■ Build the model

- A covalent bond can also break by **heterolytic fission** : BOTH electrons of the shared pair move to one atom, forming ions. A double-barbed curly arrow shows an electron PAIR moving.
- A **nucleophile** is an electron-pair donor (OH^- , CN^- , NH_3) that attacks a $\delta+$ carbon. An **electrophile** is an electron-pair acceptor (Br_2 , H^+) that accepts a pair from an electron-rich region such as a C=C double bond. In the substitution of bromoethane, OH^- donates its pair to the $\delta+$ carbon while Br^- leaves with the bonding pair: $\text{CH}_3\text{CH}_2\text{Br} + \text{OH}^- \rightarrow \text{CH}_3\text{CH}_2\text{OH} + \text{Br}^-$.
- heterolytic fission: Bond breaking where one atom takes both shared electrons; nucleophile: An electron-pair donor attracted to a positive or $\delta+$ centre; electrophile: An electron-pair acceptor attracted to an electron-rich region.

Heterolytic fission: the PAIR moves to one atom



Nucleophilic substitution: OH^- attacks bromoethane



Nucleophile = electron-pair donor; electrophile = electron-pair acceptor

Double-barbed arrows move an electron PAIR

Electron-pair sharing reactions: original worked-case diagram

■ Choose evidence that can test it

- The reaction rate depends on the concentrations of BOTH bromoethane and hydroxide, supporting a mechanism in which both meet in the key step. Bromide ions appear in solution as the alcohol forms, matching the leaving-group prediction.
- Bromine water loses its orange colour with an alkene but not with the corresponding alkane – evidence that the electron-rich $\text{C}=\text{C}$ is what attracts the electrophile.

■ Work from known quantities

- Known: $\text{Mr}(\text{CH}_3\text{CH}_2\text{OH}) = 46$; $\text{Mr}(\text{CH}_3\text{CH}_2\text{Br}) = 109$; $\text{Mr}(\text{OH}^-) = 17$. Atom economy measures how much reactant mass ends in the desired product: $\text{atom economy} = \frac{\text{Mr}(\text{desired product})}{\sum \text{Mr}(\text{all reactants})} \times 100\%$. Substitution: $\frac{46}{(109 + 17)} \times 100\% = \frac{46}{126} \times 100\% = 36.5\%$. Substitution reactions have low atom economy because the leaving group carries mass away.

Example:

Chloroethane has $\text{Mr} = 64.5$. Find the atom economy of $\text{CH}_3\text{CH}_2\text{Cl} + \text{OH}^- \rightarrow \text{CH}_3\text{CH}_2\text{OH} + \text{Cl}^-$ using the same sequence: known quantities $\rightarrow$ model $\rightarrow$ relation $\rightarrow$ substitution $\rightarrow$ unit and interpretation. (Answer: $\frac{46}{(64.5 + 17)} \times 100\% = \frac{46}{81.5} \times 100\% = 56.4\%$ – a lighter leaving group raises the atom economy, though it is still far below an addition reaction's 100%.)

■ Check the conclusion and its limits

- The arrow-pushing model shows electron flow in the key step; real rates also depend on the solvent and temperature, which this scheme does not capture. Atom economy compares routes on paper; a full process evaluation needs yield, energy and safety data too.

Warn:

A nucleophile is a negative ion, so any negative ion is a nucleophile and any positive species is an electrophile. This claim is false: the definition is about electron-PAIR donation – neutral NH_3 is a nucleophile because it carries a lone pair, and Br^- is a leaving group, not an attacking nucleophile, in this substitution.

Key:

Electron-pair sharing reactions: heterolytic fission moves both shared electrons to one atom; double-barbed arrows track electron pairs. Nucleophiles donate pairs to δ^+ centres and electrophiles accept pairs from electron-rich regions; substitution and addition mechanisms follow from that flow.

Experimental programme

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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Prerequisites: read the stated quantities and units, and use arithmetic and the model conditions below.

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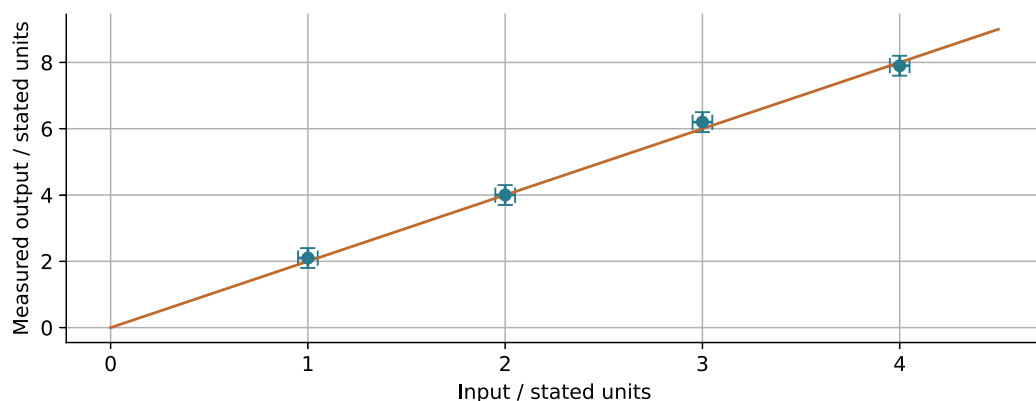
Uncertainty, gradients and model testing

■ What would explain this observation?

- A line passing near every data point is useful, but its gradient can still be uncertain. A graph is evidence for a model within the measurement range.

■ Build the model

- Random variation makes repeated readings differ.
- **Systematic error** shifts results consistently.
- Absolute **uncertainty** has the measured unit; relative or percentage uncertainty compares uncertainty with the measured value.
- uncertainty: A quantified limitation on a measured result.
- systematic error: A consistent measurement bias.



Uncertainty, gradients and model testing: original worked-case diagram

■ Choose evidence that can test it

- For a product or quotient, adding fractional uncertainties is a common maximum-uncertainty approximation.
- For a difference, add absolute uncertainties.
- A nonzero intercept can reveal an offset or an incomplete model.

In the lab: Show units on axes and choose a sensible scale. Plot uncertainty bars where justified, draw a best-fit line rather than joining every point, and estimate steepest and shallowest plausible gradients when the course method calls for them.

■ Work from known quantities

- Known: length = 50.0 mm with uncertainty 1.0 mm.
- Percentage uncertainty = absolute uncertainty/value $\times 100 = 1.0/50.0 \times 100 = 2.0\%$.
- For a quotient of two independently measured quantities with maximum percentage uncertainties 2% and 3%, the summed maximum estimate is 5%.

Your turn: A 40 cm reading has an absolute uncertainty of 1 cm. Find percentage uncertainty.

■ Check the conclusion and its limits

- Repeating readings reduces random uncertainty in a mean but does not automatically remove a zero error.
- Do not quote more decimal places than your measurement can support.

Not true: Repeating a measurement always removes a calibration offset. Repeating and averaging reduces RANDOM scatter in the mean. A calibration offset shifts every reading identically, so repeats land on the same wrong value; only recalibration removes it.

Key:

- Random variation makes repeated readings differ.
- For a product or quotient, adding fractional uncertainties is a common maximum-uncertainty approximation.
- Watch for: repeating readings reduces random uncertainty in a mean but does not automatically remove a zero error.

■ Write from a percentage uncertainty of 8 percent

- A reading is 25 cm and the absolute uncertainty is 2 cm.
- Find the percentage uncertainty. Repeating the measurement does not remove a calibration offset.