

Supported Focus Companion

International Baccalaureate · IB Diploma · Chemistry · HL

Source-supported focus teaching; remaining qualification requirements are held

;

Original teaching and practice · exercises.lol

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Introduction to the particulate nature of matter · supported focus

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

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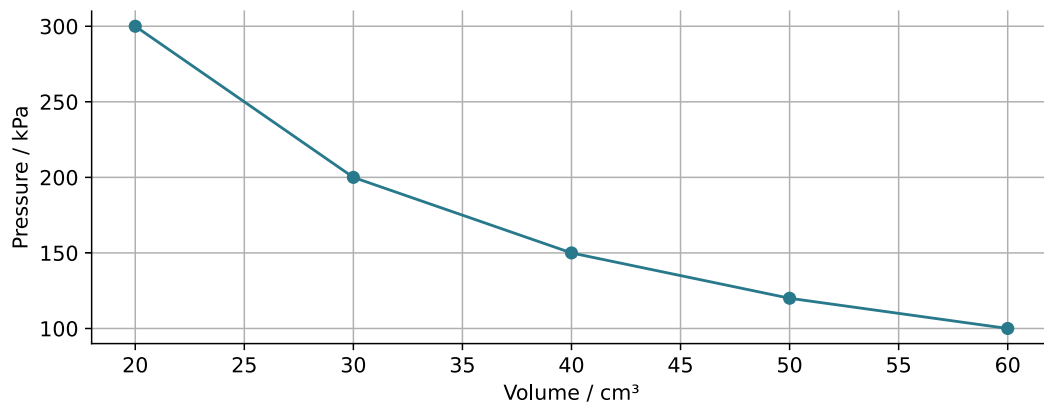
Gas models and absolute temperature

■ 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**.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

At fixed volume, pressure is 120 kPa at 300 K. Find pressure at 350 K. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

A gas pressure ratio can always use Celsius temperatures. This claim is false: 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.

Key:

Gas models and absolute temperature: 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.

The nuclear atom · supported focus

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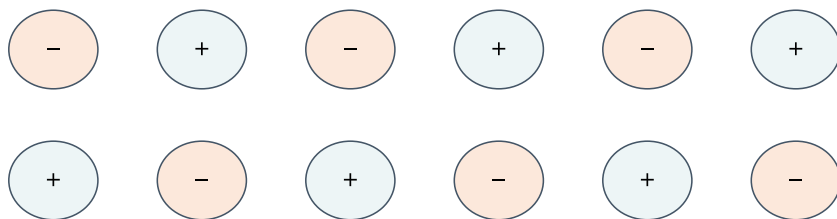
Particles, bonding and bulk properties

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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

■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

An atom has atomic number 17 and mass number 35. Find its number of neutrons. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

All covalent substances have low melting points. This claim is false: Melting a simple molecular substance usually overcomes intermolecular attractions; it does not require breaking all covalent bonds within each molecule.

Key:

Particles, bonding and bulk properties: 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.

Counting particles by mass: the mole · supported focus

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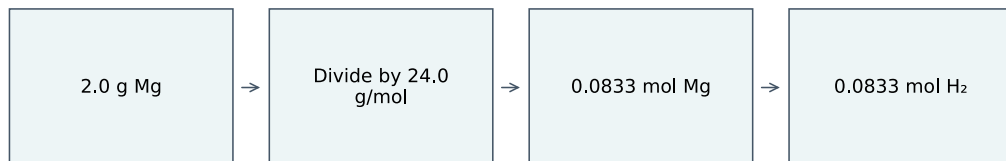
Amounts, equations and limiting reagents

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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 H₂ = amount Mg = 0.0833 mol.

Example:

Calculate amount in 5.0 g of a substance with molar mass 100 g per mole. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

Balanced equation coefficients always give reactant mass ratios. This claim is false: Excess acid means acid does not limit the stated calculation. A coefficient of 2 before HCl does not double the hydrogen amount.

Key:

Amounts, equations and limiting reagents: 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.

Functional groups: classification of organic compounds · supported focus

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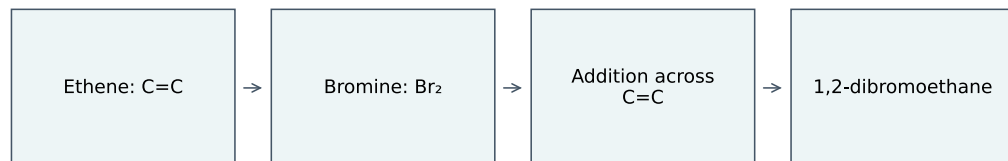
Organic structures and reaction pathways

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

How many moles of bromine react completely with 0.15 mol ethene? Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

Compounds with the same molecular formula always have identical properties. This claim is false: Bromine decolourization provides evidence of unsaturation in an appropriate test. It is not proof that an unknown sample is specifically ethene.

Key:

Organic structures and reaction pathways: 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.

Measuring enthalpy change · supported focus

International Baccalaureate · IB Diploma · Chemistry · HL

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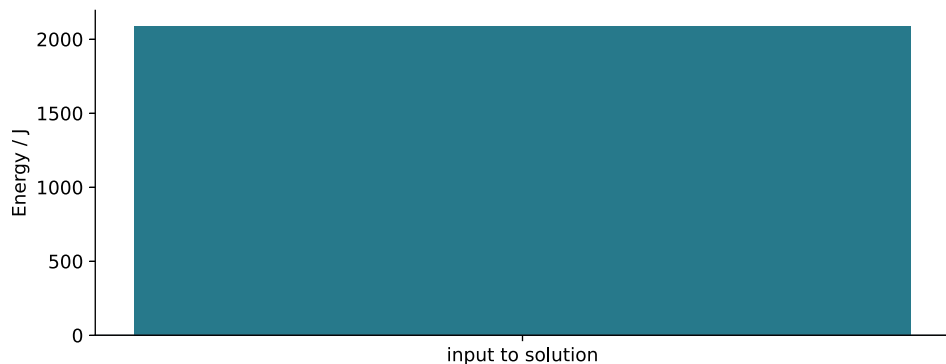
Calorimetry and chemical energy

■ 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**.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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, $\Delta H = -q/n = -2.09/0.050 = -41.8 \text{ kJ per mole}$.

Example:

50 g water rises 4 °C. Use $c = 4.2 \text{ J per gram per degree}$ to find q . Use the same sequence: known quantities $\rightarrow$ model $\rightarrow$ relation $\rightarrow$ substitution $\rightarrow$ unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

Bond breaking releases energy. This claim is false: Heat loss usually lowers the observed temperature rise. The solution gaining heat and the reaction losing heat have opposite signs.

Key:

Calorimetry and chemical energy: 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.

Energy from fuels · supported focus

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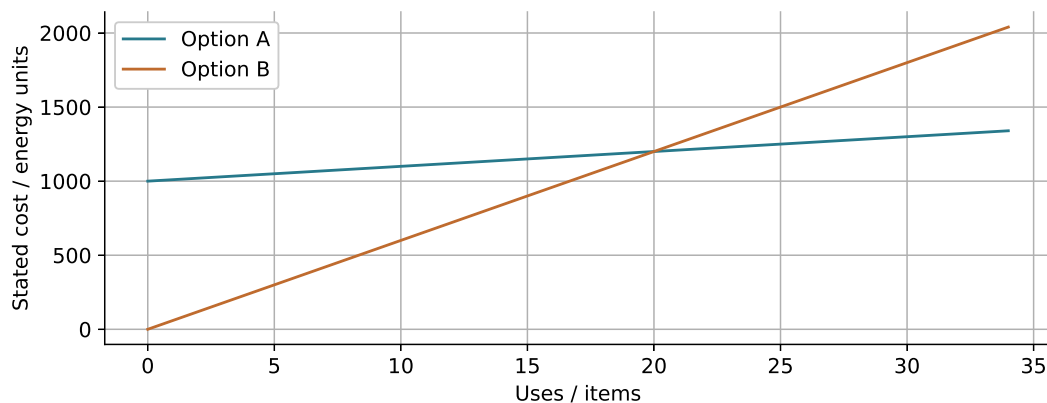
Atmosphere, resources and life-cycle decisions

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

A reusable item costs 600 energy units initially and 5 per use; disposables cost 35 per use. Find break-even uses. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

A recyclable product is always recycled after use. This claim is false: A break-even model is sensitive to cleaning, transport and disposal assumptions. Recyclable does not guarantee that an item is actually collected and recycled.

Key:

Atmosphere, resources and life-cycle decisions: 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.

Entropy and spontaneity · supported focus

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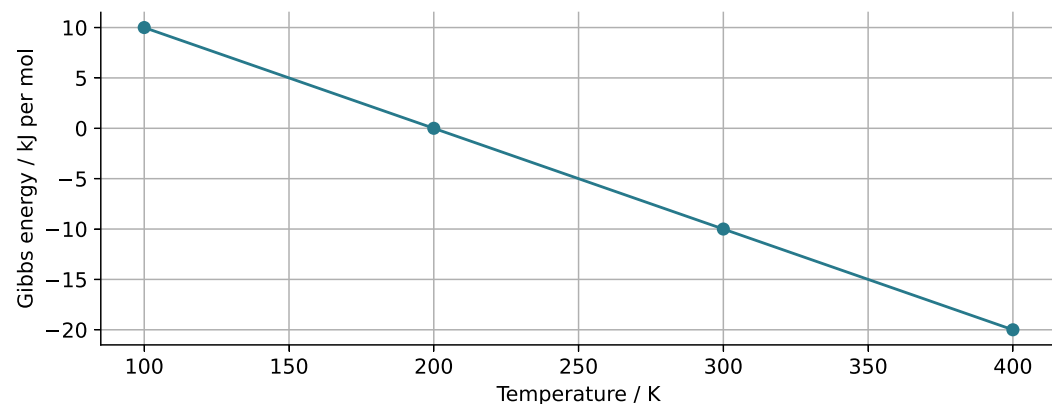
Entropy , Gibbs energy and feasibility

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

$\Delta H = 30$ kJ per mole, $\Delta S = 0.10$ kJ per kelvin per mole and $T = 400$ K. Find ΔG . Use the same sequence: known quantities $\rightarrow$ model $\rightarrow$ relation $\rightarrow$ substitution $\rightarrow$ unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

A thermodynamically favourable reaction must be fast. This claim is false: 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.

Key:

Entropy, Gibbs energy and feasibility: 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.

How much? The amount of chemical change · supported focus

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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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.
- 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 endpoint, swirl, and read the meniscus at eye level. Use a white tile and appropriate eye protection.

■ Work from known quantities

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

Burette readings are 1.40 and 23.65 cubic centimetres. Calculate titre. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

A flask must be rinsed with analyte to avoid dilution of its mole amount. This claim is false: 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.

Key:

Titration and a defensible concentration: 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.

How fast? The rate of chemical change · supported focus

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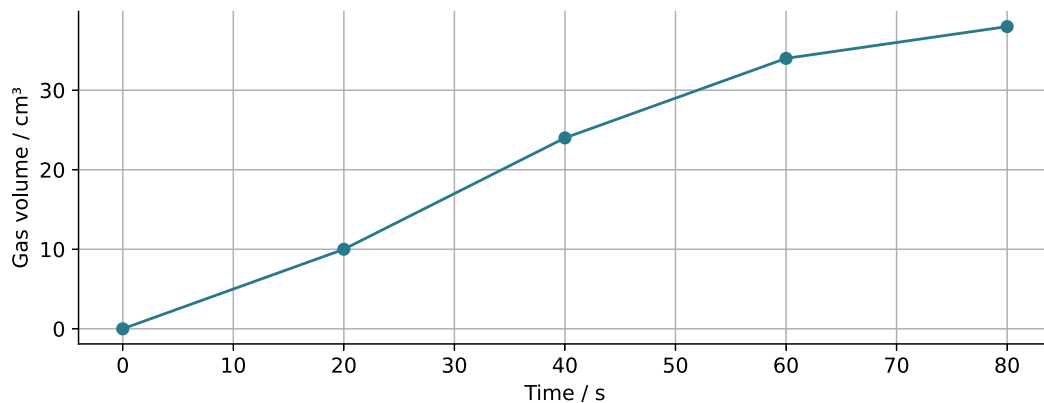
Rates, catalysts and reliable endpoints

■ What would explain this observation?

- A faster reaction finishes sooner, but it need not make more product. **Rate** and final yield answer different questions.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

A reaction makes 24 cubic centimetres of gas in 40 s. Find average rate. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

A faster reaction always produces a larger final amount of product. This claim is false: 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.

Key:

Rates, catalysts and reliable endpoints: 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.

How far? The extent of chemical change · supported focus

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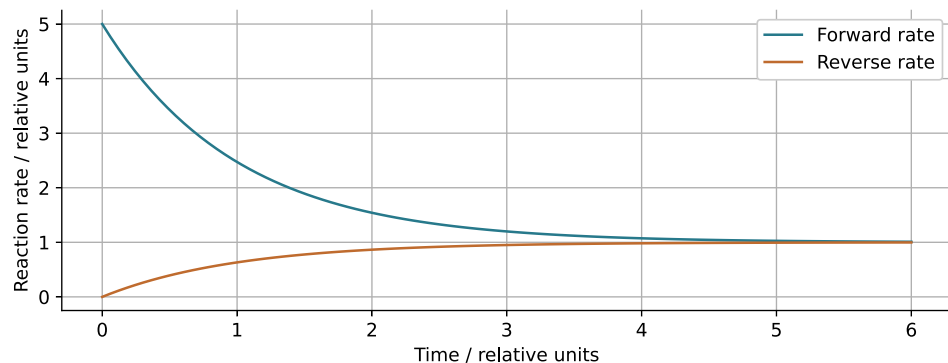
Equilibrium and changing conditions

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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 $K = 1$.

Example:

For $A \rightleftharpoons B$, calculate $[B]/[A]$ when $[B]=0.8$ and $[A]=0.2$. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

A catalyst changes the equilibrium constant at a fixed temperature. This claim is false: Do not use a catalyst to claim a larger equilibrium yield. For heterogeneous equilibria, pure solids are omitted from the usual equilibrium expression.

Key:

Equilibrium and changing conditions: 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.

Electron transfer reactions · supported focus

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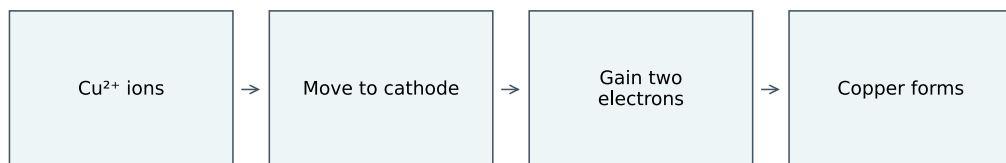
Redox and electrolysis

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- 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.

Example:

How many moles of electrons reduce 0.20 mol of Cu^{2+} to copper? Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

Reduction means loss of electrons. This claim is false: 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.

Key:

Redox and electrolysis: 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.

Experimental programme · supported focus

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

Supported HL focus. First assessment 2025; current brief acquired; full Chemistry guide not acquired. Remaining guide, assessment and practical requirements retain their recorded holds.

Prerequisites: read the stated quantities and units, use arithmetic and the model conditions below. Each lesson develops its own method before independent transfer.

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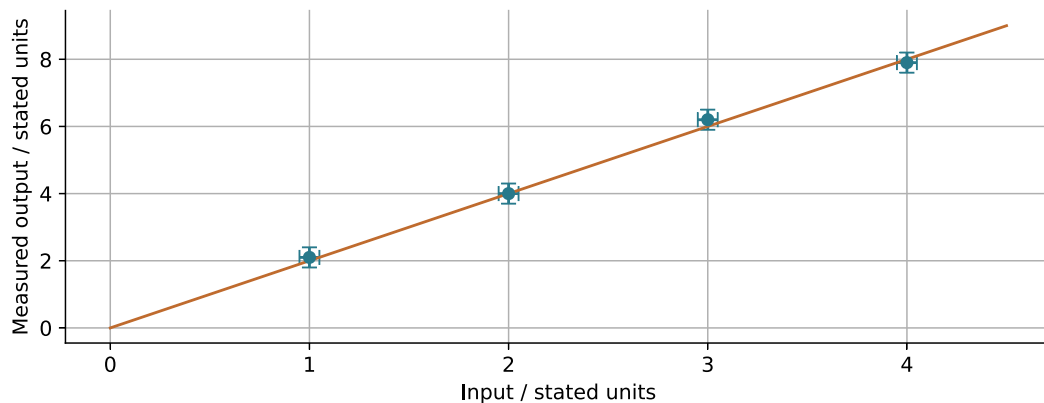
Uncertainty, gradients and model testing

■ 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.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ 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.



■ 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.
- 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

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- Known: length = 50.0 mm with uncertainty 1.0 mm. Percentage uncertainty = $\text{absolute uncertainty} / \text{value} \times 100 = 1.0 / 50.0 \times 100 = 2.0\%$. For a quotient of two independently measured quantities with maximum percentage uncertainties 2% and 3%, the summed maximum estimate is 5%.

Example:

A 40 cm reading has an absolute uncertainty of 1 cm. Find percentage uncertainty. Use the same sequence: known quantities → model → relation → substitution → unit and interpretation.

■ 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.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

Repeating a measurement always removes a calibration offset. This claim is false: 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.

Key:

Uncertainty, gradients and model testing: 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.