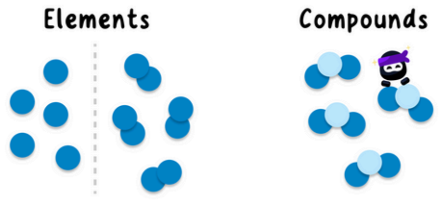
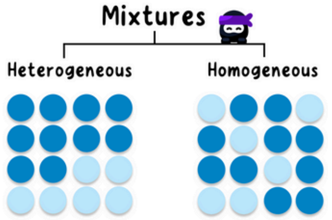
IB CHEMISTRY S1.1 INTRODUCTION TO THE PARTICULATE NATURE OF MATTER

Properties of elements, compounds, and mixtures

- **Element**: pure substance made of **one type** of atom; cannot be chemically broken down
- **Compound**: pure substance of **two or more** elements chemically bonded in a **fixed**, **simple whole-number ratio**; properties differ from constituent elements
- **Mixture**: physical combination of **two or more** substances (elements and/or compounds) in **no fixed ratio**; components retain their properties and can be separated by physical means
    - **Homogeneous mixture**: uniform composition throughout(e.g., air, salt solution); single visible phase
    - **Heterogeneous mixture**: non-uniform composition(e.g., oil+water, granite); two or more visible phases





Kinetic molecular theory

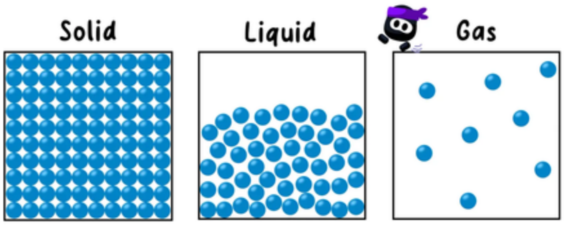
- Particles are in constant random motion: $Ek \propto T$
- Collisions are elastic(on average energy conserved)
- Intermolecular forces determine proximity and order

Exam Tip!

If a question says “explain using KMT”,mention particle spacing, motion, intermolecular forces, and how these account for observed property.

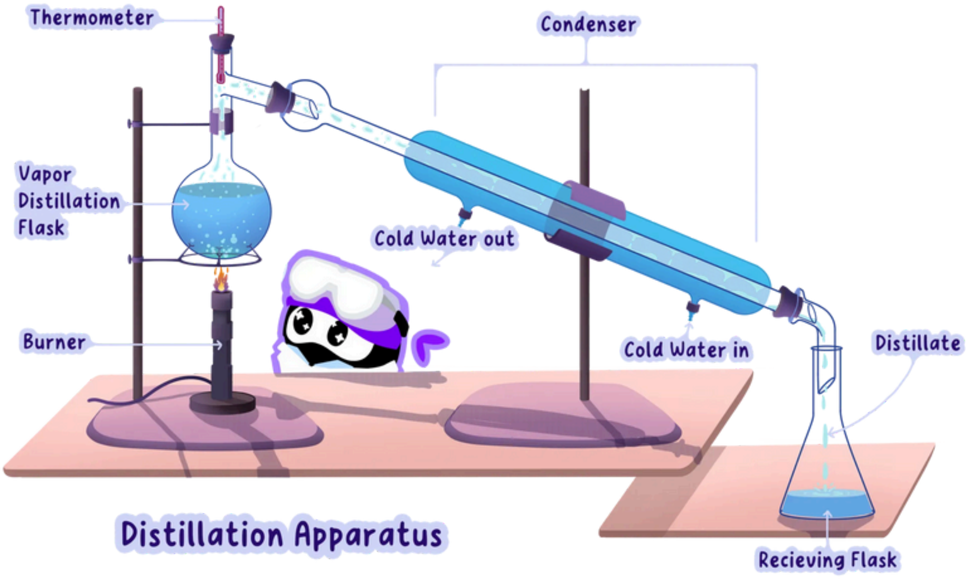
State of matter and state symbols

- **Solid (s)**: particles closely packed in fixed positions (vibrate only); strong forces; fixed shape & volume; **lowest Ek**
- **Liquid (l)**: particles close but able to slide past; moderate forces; fixed volume, variable shape; **medium Ek**
- **Gas (g)**: particles far apart, negligible forces (ideal model); variable volume & shape; compressible; **highest Ek**
- **Aqueous (aq):** dissolved in water (not a state of matter!)



Separation techniques you must know

- **Solvation (dissolving)**: solute particles become surrounded by solvent particles; used to prepare solutions.
- **Filtration**: separates insoluble solid from liquid using a porous barrier(e.g., sand from water)
- **Recrystallization**: purifies a solid by dissolving hot then crystallizing on cooling; impurities remain in mother liquor
- **Evaporation**: removes solvent to obtain dissolved solid; gentle heating increases rate
- **Distillation**: separates miscible liquids by different boiling points; simple(distill solvent from solute) vs fractional(separate close b.p. liquids)
- **Paper chromatography**: separates mixture based on different affinities between mobile phase(solvent) and stationary phase(paper).



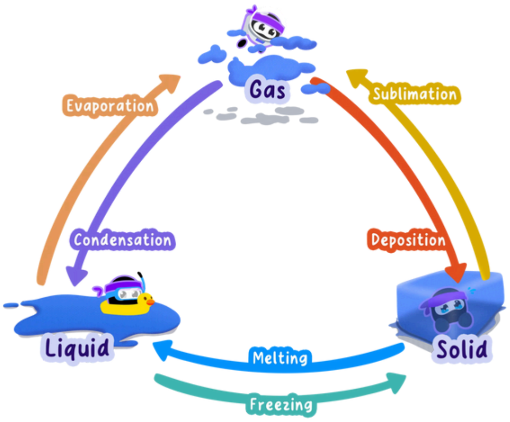
Distillation Apparatus

Example:

You can separate ethanol (boiling point ~78°C) from water (boiling point ~100°C) using a distillation apparatus.

Changes of state

- **Melting**: solid → liquid (endothermic)
- **Freezing**: liquid → solid (exothermic)
- **Vaporization**: liquid → gas (endothermic); includes:
    - **Evaporation**: only at the **surface of the liquid**, at any temperature, when particles near the surface absorb enough energy to escape
    - **Boiling**: throughout the **entire liquid** at a specific temperature where particles throughout the liquid gain enough energy to break free
- **Condensation**: gas → liquid (exothermic)
- **Sublimation**: solid → gas (endothermic)
- **Deposition**: gas → solid (exothermic)



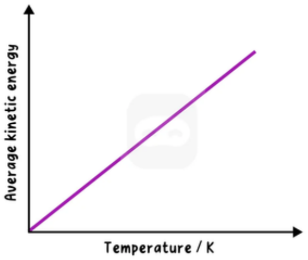
Common Question

What is the change of state for a solid to a gas?

A. Vaporization
B. Deposition
C. Condensation
D. Sublimation

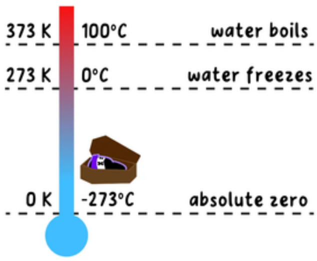
Temperature and kinetic energy

- Absolute temperature T(K) is **directly proportional** to average kinetic energy Ek of particles: **Ek ∝ T**
- At the same T, all gases have the same average Ek (regardless of identity); **lighter** molecules move **faster** on average



Kelvin vs Celsius

- Kelvin is SI base unit
- Increment size identical to °C, but different zero points
- Conversion formulas you must use exactly
    - T(K) = T(°C) + 273.15
    - T(°C) = T(K) − 273.15

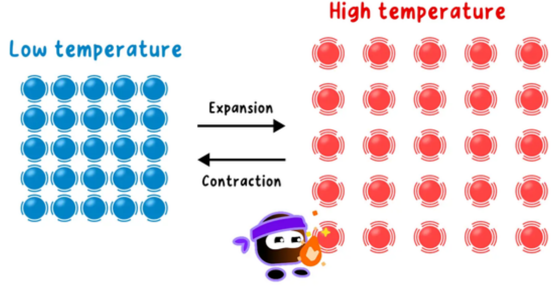


Example:

- 25°C → K: T = 25 + 273.15 = 298.15 K (≈298 K)
- -10°C → K: T = −10 + 273.15 = 263.15 K
- 310K → °C: T = 310 − 273.15 = 36.85°C

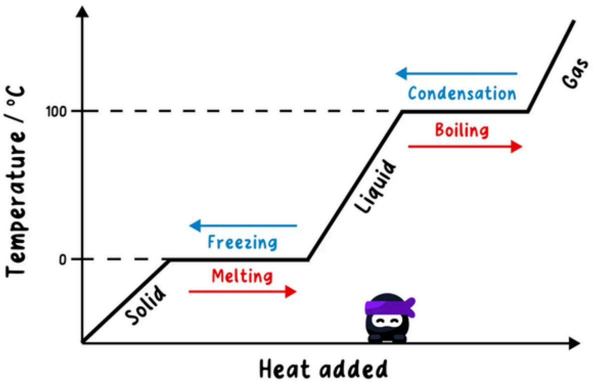
Interpreting changes with temperatures and during phase changes

- Temperature **rises** → particles gain kinetic energy → move more vigorously thus occupy more space → expansion
- Temperature **falls** → particles lose kinetic energy → slow down thus take up less space → contraction
- **Phase change** = **no** temperature change = transferred energy is used to break or form intermolecular bonds



Heating curve

- **Flat regions** = phase changes where temperature remains constant while energy is used to break or form bonds.
- **Sloped regions** = temperature changes as kinetic energy increases or decreases.



Exam Tip!

If asked why temperature is constant during melting/boiling, explicitly state:

- energy input changes potential energy to overcome intermolecular forces, not kinetic energy,
- hence temperature remains constant.

IB CHEMISTRY S1.2
THE NUCLEAR ATOM

Atomic structure, nuclear symbol, subatomic particles

Atoms consist of:

- **Protons**: positive charge (+1), relative mass = 1
- **Neutrons**: no charge (0), relative mass = 1
- **Electrons**: negative charge (−1), relative mass ≈ 0 (negligible compared to nucleons)

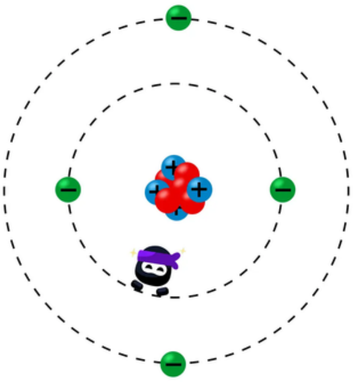
In the **nuclear model** of an atom:

- **Nucleus**: dense, contains nearly all the mass; positively charged because of protons
- **Electrons**: occupy surrounding space, organized into energy levels

Nuclear symbol format:

$$^{A}_{Z}X^{q}$$

- **X** = element symbol
- **A** = mass number = protons + neutrons
- **Z** = atomic number = number of protons
- **q** = charge on ion (omitted if neutral)



Isotopes and calculations with relative atomic masses

- **Isotopes** = atoms of the same element (same Z, different A) → **same** # protons and electrons, **different** # neutrons
- **Chemical properties**: identical(because determined by electron configuration)
- **Physical properties**: may differ (density, mass, rate of diffusion, melting/boiling points slightly)

- **Relative atomic mass (Ar)** = weighted average of isotopes compared to 1/12 of mass of carbon-12 atom

$$A_r = \frac{\sum(\%\text{ abundance } \times \text{ isotopic mass })}{100}$$

Example:

For **magnesium** isotopes: 79% Mg-24, 10% Mg-25 and 11% Mg-26:

Ar = (24 × 79 + 25 × 10 + 26 × 11) / 100 = = 24.32

Common Question

What is the relative atomic mass of a sample of bromine containing 55% of the Br-79 isotope and 45% of the Br-81 isotope?

A. 79.8
B. 79.9
C. 80.0
D. 80.2

Mass spectrometry and isotopic composition (HL)

- **Mass spectrometry**: technique to separate isotopes by mass-to-charge ratio (m/z)
- **Steps** (not assessed in detail): ionization, acceleration, deflection, detection
- **Key output**: mass spectrum = peaks showing isotopes of an element

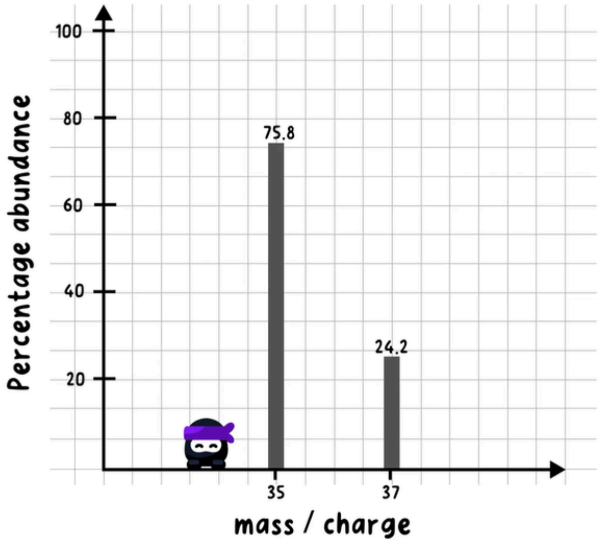
Interpreting spectra:

- **x-axis** = m/z (≈mass number A for singly charged ions)
- **y-axis** = relative abundance (%)
- **Multiple peaks** = multiple isotopes present
- **Relative atomic mass** calculated using weighted average of isotopes from the spectrum

Exam Tip!

If given mass spectrum, explicitly multiply mass × relative abundance, sum, then divide by total abundance.

Example:



In the graph above:

- **Two** main peaks: m/z = 35 and m/z = 37 ← represent naturally occuring isotopes
- **Cl-35**: higher peak = greater abundance
- **Cl-37**: lower peak = lesser abundance

Ar = (75.8 × 35 + 24.2 × 37) / 100 = 35.45

IB CHEMISTRY S1.3
ELECTRON CONFIGURATIONS

Atomic structure, uclear symbol, subatomic particles

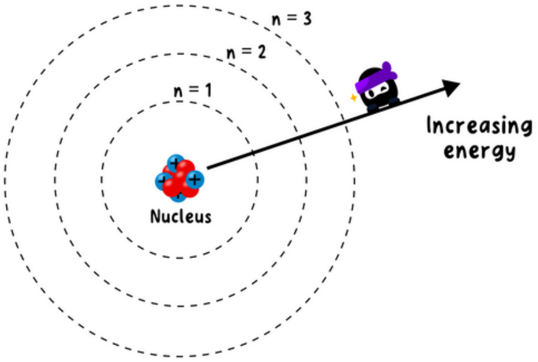
- **Emission spectrum** forms when excited electrons drop to lower energy levels and emit photons with **energy $E = hf = hc / \lambda$**
- Qualitatively:
  - **shorter λ** → higher f → higher E
  - **longer λ** → lower f → lower E

- **Continuous spectrum**: all λ present with no gaps (e.g., white light through prism)

- **Absorption spectrum**: discrete dark lines at specific λ only against a continuous background.

Atomic structure, uclear symbol, subatomic particles

Main energy level n can hold max **$2n^2$** electrons:
n = 1 → 2 **n = 2** → 8 **n = 3** → 18 **n = 4** → 32

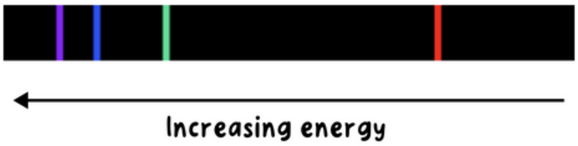


Note!

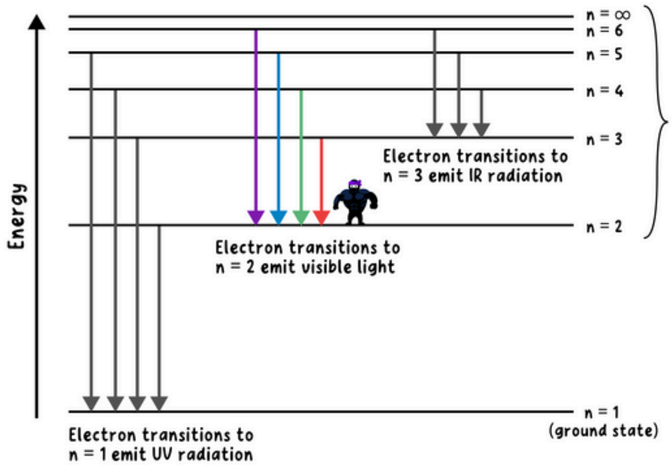
Within a real atom, sublevel energies cause the observed filling order; $2n^2$ is a capacity, not a filling sequence

Hydrogen line spectrum & discrete energy levels

- **Hydrogen in emission** shows distinct lines because electrons move between quantized energy levels; lines get closer (**converge**) at **higher energies**

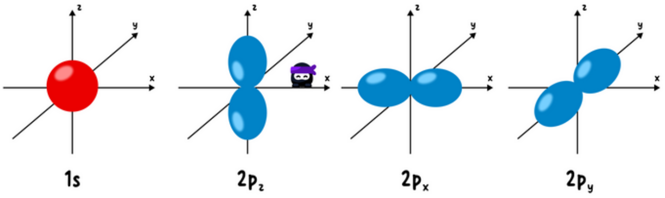


- **Transitions**:
  - ending at **n = 1** produce high-energy UV lines;
  - ending at **n = 2** produce visible lines;
  - ending at **n = 3** produce IR lines.
- **Convergence** means levels get closer together as n increases → approaches the **ionization limit**



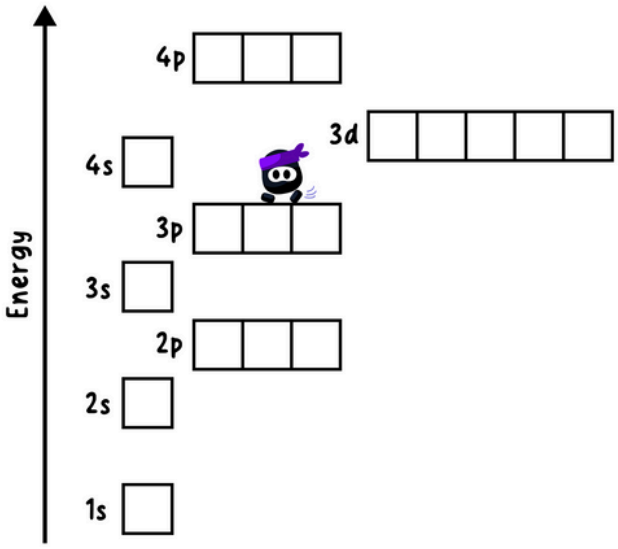
Sublevels and orbital shapes

- Main level splits into **sublevels** of increasing energy: s < p < d < f
- Orbitals are regions of **high electron probability**; **each** orbital holds **max 2** electrons
- Counts **per sublevel**: 1 s orbital (2e⁻); 3 p orbitals (6e⁻); 5 d orbitals (10e⁻); 7 f orbitals (14e⁻)
- Recognize **shapes**:
  - **s orbital**: spherical, centered on nucleus
  - **p orbitals**: three dumbbells oriented along x, y, z (denoted p_x, py, pz)
  - **d and f orbitals**: beyond IB scope



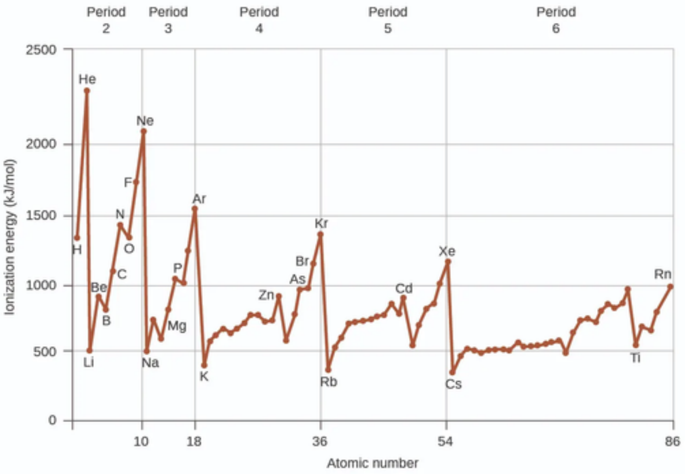
Aufbau, Hund, Pauli; configurations up to Z = 36, Cr & Cu exceptions

- **Filling order** (**Aufbau**): 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s…(use diagonal/energy ordering)
- **Hund's rule**: within a sublevel, place 1e⁻ per orbital with parallel spins before pairing
- **Pauli exclusion**: an orbital holds max 2e⁻ with opposite spins (↑↓)
- Orbital diagrams:
  - **boxes** = orbitals
  - **arrows** = electrons with spin
  - fill per Hund then pair
- Canonical **exceptions**:
  - Cr: $[Ar]3d^5 4s^1$ (not $[Ar]3d^4 4s^2$) → **stabilized half-filled d**
  - Cu: $[Ar]3d^{10} 4s^1$ (not $[Ar]3d^9 4s^2$) → **stabilized filled d**
- For **ions**, adjust from the outermost shells first:
  - **Cations** of transition metals lose **4s before 3d**:
    Fe $[Ar]3d^6 4s^2$ → Fe^{2+} $[Ar]3d^6$; Fe^{3+} → $[Ar]3d^5$
  - **Anions** add to p sublevel until noble gas: Cl^- → [Ar]



Covergence limit, first IE trends & calculations (HL)

- **Convergence limit**: highest-frequency line as levels converge corresponds to removing e^- from ground state to $n = \infty$ (ionization)
- **First ionization energy** (IE_1): $X(g) \rightarrow X^+(g) + e^-$; energy required per mole
- **Trends across a period**: IE generally increases (↑effective nuclear charge, ↓atomic radius)
- **Discontinuities**:
    - ***Group 2 → 13*** (Be → B, Mg → Al) :drop because the e^- removed is from higher-energy p vs s sublevel
    - ***Group 15 → 16*** (N → O, P → S): drop due to electron repulsion when pairing in a p orbital (half-filled p^3 is relatively stable)
- **Trend down a group**: IE decreases(↑n, ↑shielding, ↑radius) despite higher nuclear charge
- Calculate IE from convergence limit:
    - **Per atom**: $\Delta E = hf = hc/\lambda$
    - **Per mole**: $IE = N_A \cdot hc/\lambda$ (or $N_A \cdot h \cdot f$); give in J mol^{-1} or kJ mol^{-1}



Source:https://pressbooks.online.ucf.edu/chemistryfundamentals/chapter/8-8-electron-affinities-and-metallic-character-chemistry-libretexts/

Common Mistake

- Don't mistakenly assume that ionization energy always increases across a period.
- Always consider **sublevel stability** and **electron repulsion effects** for discontinuities.

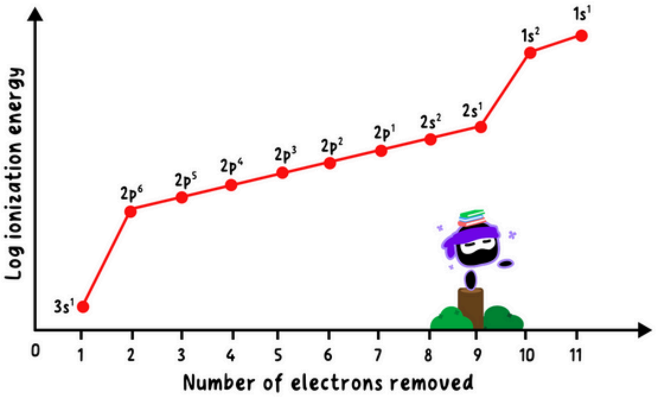
Succesive IEs & deducing group (HL)

- **Successive IE** data show **large "jump"** after all valence e^- are removed; **position of jump** → number of valence electrons → **group**

Example:

Consider the following dataset (kJ mol^{-1}):

- Element X:
    - $IE_1 = 738$,
    - $IE_2 = 1451$,
    - $IE_3 = 7733$,
    - $IE_4 = 10500$...
- Big jump between **IE_2** and **IE_3** → **two** valence e^- → **Group 2**



This graph shows successive ionization energies of **sodium**, starting for the 3rd **energy level**.

IB CHEMISTRY S1.4 COUNTING PARTICLES BY MASS: THE MOLE

The mole and Avogadro's constant; counting entities

- n (mol) = **amount of substance**
- **1 mol** contains NA elementary entities exactly
- **Elementary entity** can be atom/molecule/ion/electron/other specified group
- **Relationships**: $N = n \cdot NA$ and $n = N/NA$
- **Units**: NA in mol^{-1}; n in mol; N is a pure count(no units)
- Use the **booklet** value for NA in calculations

Exam Tip!

If the question asks for "number of oxygen atoms in x mol of CO_2", first find molecules (N), then multiply by 2.

Relative atomic and formula masses

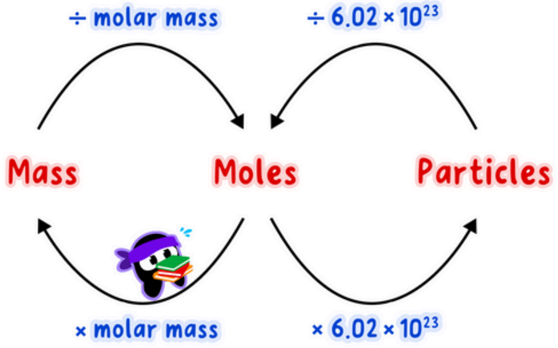
- **Ar** is average mass of an atom on scale where $Ar(^{12}C) = 12$ exactly
- **Mr** is sum of Ar of all atoms in a formula unit
- Ar and Mr have **no units**
- Use Ar from the **data booklet** (to 2 d.p.) for all calculations
- For **molecules**(covalent) use "relative **molecular** mass"; for ionic/formula units use "relative **formula** mass"

Example:

$Mr\ ((NH_4)_2SO_4) =$
$= 2(14.01+4\cdot1.01)+32.07+4\cdot16.00 = 132.13$

Molar mass

- **Molar mass M** is in $g\ mol^{-1}$ and is numerically ≈ Mr for a substance
- **Core equations**:
  - $n=m/M$ → moles
  - $m=n\cdot M$ → mass
  - $N=n\cdot NA$ → particles



÷ molar mass

÷ 6.02×10^{23}

Mass

Moles

Particles

× molar mass

× 6.02×10^{23}

Empirical vs molecular formula; % composition

- **Empirical formula** (EF) = simplest whole-number ratio
- **Molecular formula** (MF) = actual numbers in a molecule
- **% to EF steps**:
  - Assume 100 g
  - Convert % to g
  - Covert g to mol
  - Divide by smallest
  - Scale to whole numbers
- **MF from EF**:
  - Find n= M(MF)/M(EF)
  - Multiply EF subscripts by n
  - If ratio close to x.50, multiply all by 2; for x.33/x.66 multiply by 3, etc.

Example:

For 40.00% **C**, 6.71% **H**, 53.29% **O**:

- Moles: C = 3.33, H = 6.66, O = 3.33
- Ratio: 3.33:6.66:3.33 = 1:2:1
- EF = CH_2O
- If M = 180 $g\ mol^{-1}$ then n = 180/30 = 6 → $MF=C_6H_{12}O_6$

Molar concentration and solutions calculations

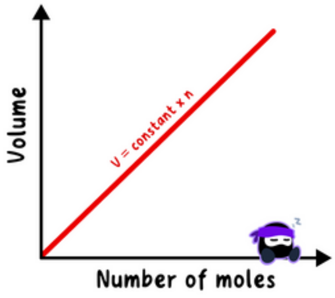
- **Concentration** notation: [X] = c in $mol\ dm^{-3}$; also use $g\ dm^{-3}$ when required
- Key relations:
  - $n = C\cdot V$ (with V in dm^3)
  - $C = n/V$
  - $C\ (g\ dm^{-3}) = m/V$
  - Link mass via $m = C\cdot V\cdot M$
- **Unit** facts: 1 dm^3 = 1 L; 1 cm^3 = 1 mL = 10^{-3} dm^3

Exam Tip!

If a **dilution** is involved, use $\mathbf{C_1V_1 = C_2V_2}$ (same solute moles before = after) and then convert to requested units

Avogadro's law and gas volumes

- **Avogadro's law**: at same T and p, equal volumes of gases contain equal numbers of molecules → **V ∝ n**
- The **molar volume** you should know by heart: STP (273 K, 1.00×10^5 Pa) ≈ 22.7 $dm^3\ mol^{-1}$ (the value can be also found in the **data booklet**)
- In gas **stoichiometry** under **same conditions**, volume ratios = mole ratios



IB CHEMISTRY
S1.5 IDEAL GASES

Ideal gas model & assumptions

- **Ideal gas** =
  - point particles in **constant random** motion
  - **negligible** particle volume
  - **no** intermolecular forces
  - **perfectly elastic** collisions
  - **continuous**, **random** straight-line motion between collisions
- **Temperature** relates to average kinetic energy; collision duration ≪ time between collisions
- **Pressure** arises from particle collisions with container walls; time-averaged, isotropic
- **Internal energy** depends only on temperature (not on volume or pressure)

Common Mistake

Don't say that gas has "negligible mass"; it's negligible **volume of particles**, not mass.

Common Question

Suggest why ammonia vapour (NH_3) deviates significantly from ideal behaviour when the gas is cooled, whereas neon (Ne) does not. [2 marks]

Real gases & limitations

- **Real gases deviate** notably at **low T** (IMFs become significant → condensation) and **high p** (finite molecular volume matters)
- **Deviation sources:**
  - attractive forces (lower p than ideal)
  - repulsive/finite size (raises p vs ideal at very high p)
  - possible association (e.g., H-bonding) and non-sphericity
- Conditions **closest to ideal**: high T, low p, small/nonpolar molecules (He, Ne, H_2, N_2)

Note!

No van der Waals math required here: just the **qualitative reasons** (IMFs and finite size).

Ideal gas equation & combined gas law

- **Core equations** (SI units only):
  - $PV = nRT$
  - $P_1V_1/T_1 = P_2V_2/T_2$
- **Units**: P in Pa (N m^{-2}), V in m^3, T in K, n in mol, $R = 8.31\ J\ mol^{-1}\ K^{-1}$
- **Conversions**: 1 atm = 1.01325×10^5 Pa; 1 dm^3 = 1 L = $1 \times 10^{-3}\ m^3$; 1 cm^3 = 1 mL = $1 \times 10^{-6}\ m^3$

Molar volume & T-p-V relationships, given fixed amount of gas

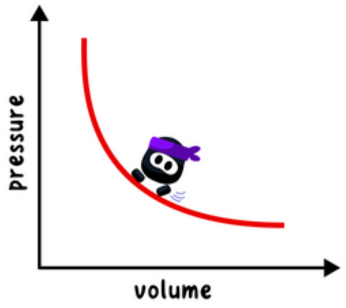
- **Molar volume Vm** is **constant** at a specified T and p (data booklet): at STP (273 K, 1.00×10^5 Pa) **Vm ≈ 22.7 $dm^3\ mol^{-1}$**

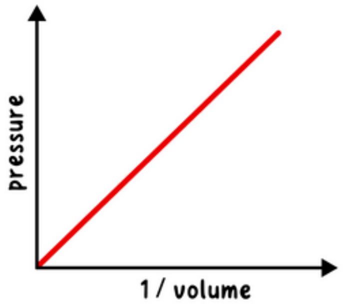
For **fixed n**:

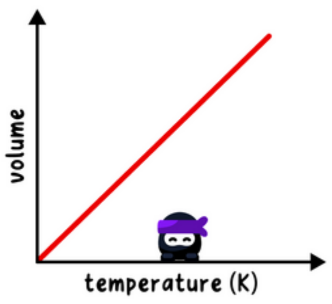
- **P-V** (at **constant T**):
  - inverse relationship
  - graph is a rectangular hyperbola
  - P·V = constant
- **V-T** (at **constant p**):
  - directly proportional
  - straight line through origin when T in K
- **P-T** (at **constant V**):
  - directly proportional
  - straight line through origin when T in K

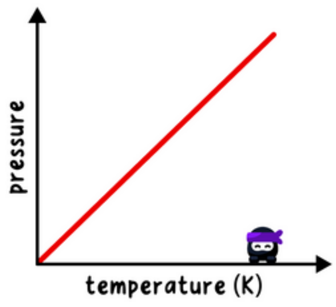
Common Mistake

Plotting versus °C gives incorrect lines (not through origin). Please don't forget to always convert to K, especially in calculations.









IB CHEMISTRY S2.1
THE IONIC MODEL

Cations, anions & predicting ionic charge

- **Ion formation**:
    - metals **lose** e⁻ → **cations** (positive)
    - non-metals **gain** e⁻ → **anions** (negative)
- Predicting charge from electron configuration: atoms tend to achieve **nearest noble-gas configuration** (minimize energy)
- Main-group shortcuts:
    - Group 1 → +1
    - Group 2 → +2
    - Group 13 → +3
    - Group 15 → −3
    - Group 16 → −2
    - Group 17 → −1
    - Group 18 → 0
- Use configurations:
    - Na $[Ne]3s^1 \rightarrow Na^+ [Ne]$
    - Mg $[Ne]3s^2 \rightarrow Mg^{2+} [Ne]$
    - O $[He]2s^2 2p^4 \rightarrow O^{2-} [Ne]$
    - N $[He]2s^2 2p^3 \rightarrow N^{3-} [Ne]$
- **Transition metals**:
    - **Variable charges** because 3d and 4s sublevels are close in energy
    - Remove **4s before 3d** when ionizing

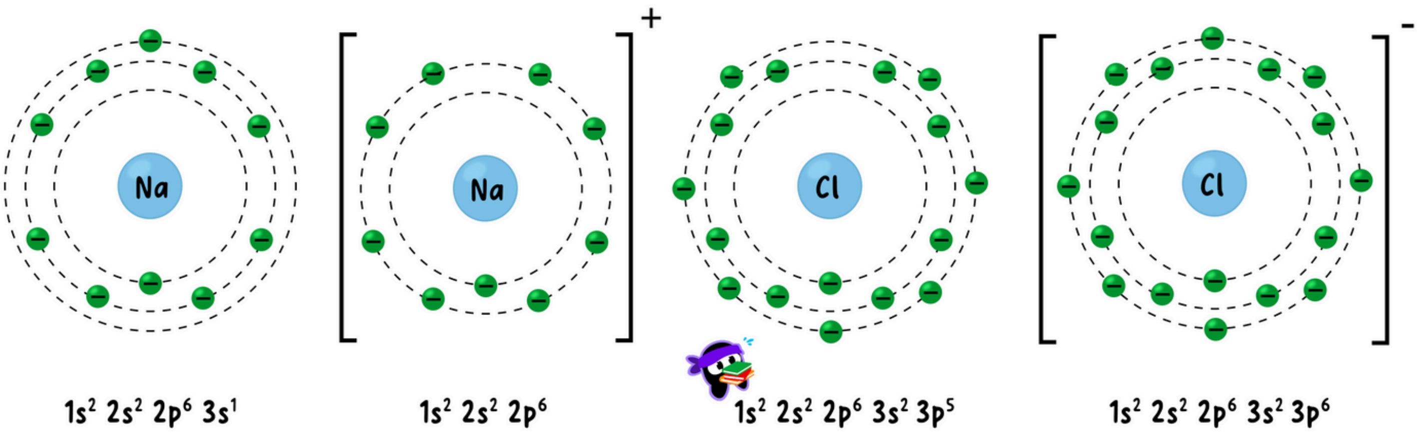
  e.g., $Fe[Ar]3d^6 4s^2 \rightarrow Fe^{2+}[Ar]3d^6$; $Fe^{3+}[Ar]3d^5$

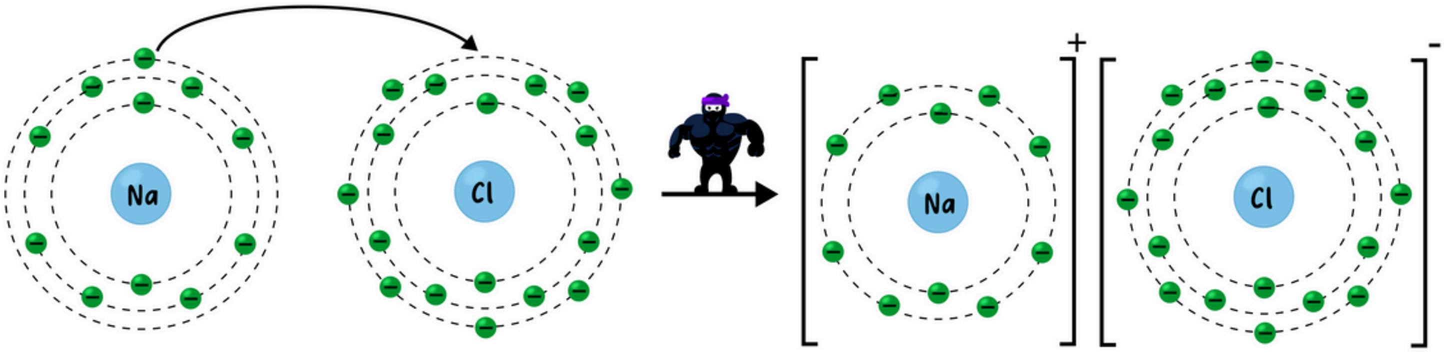
Ionic bonding: formulas & names

- **Ionic bond**:
    - Electrostatic attraction between oppositely charged ions
    - Strength **increases** with **higher charge** and ***smaller ionic radii***
- **Binary ionic naming**: cation name first, then anion with "-ide" (e.g., sodium chloride, calcium oxide); transition-metal charge in Roman numerals (e.g., copper(II) oxide)
- **Charge balance**: total positive charge = total negative charge in the empirical formula
- **Polyatomic ions** to know:
    - ammonium NH_4^+
    - hydroxide OH^-
    - nitrate NO_3^-
    - hydrogencarbonate HCO_3^-
    - carbonate CO_3^{2-}
    - sulfate SO_4^{2-}
    - phosphate PO_4^{3-}

Exam Tip!

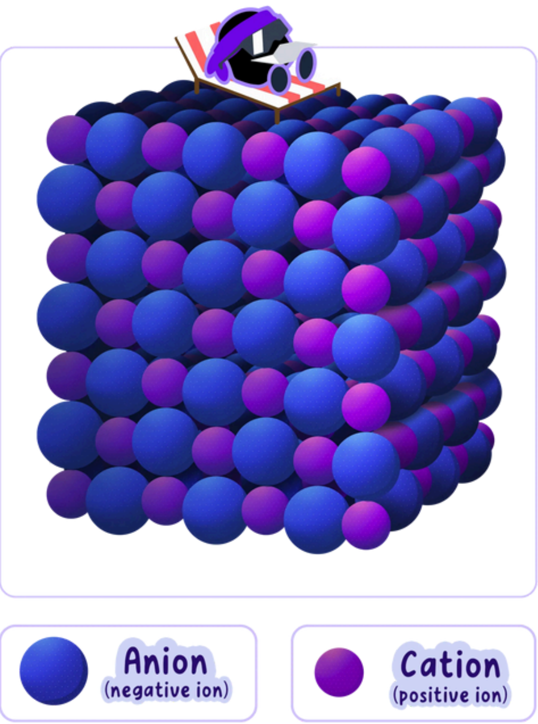
When converting **name to formula** for a transition metal, deduce the cation charge from the Roman numeral before balancing.





Ionic lattices, lattice enthalpy & properties

- **Structure**: Ionic compounds form **giant 3D lattices** of alternating cations/anions; formulas are empirical (minimum whole-number ratio)
- **Lattice enthalpy** (qualitative):
  - Energy change when 1 mol of an ionic solid forms from its gaseous ions (measure of ionic bond strength)
  - Absolute value of lattice enthalpy **increases** with **higher ionic charge** and **smaller ionic radii**
- **Volatility**: low (bonds strong → high melting/boiling points)
- **Electrical conductivity**:
  - **No** if solid (constrained ions)
  - **Yes** if molten or aqueous (free-moving ions)
- **Solubility**:
  - Often **soluble** in **polar** solvents like water (ion–dipole attractions can overcome lattice forces);
  - Trends depend on balance between lattice enthalpy and hydration enthalpy
- **Brittleness**: shattering when like charges are forced adjacent (slippage aligns same charges → repulsion)



Exam Tip!

When explaining a property, link directly to **electrostatic attractions** and, where relevant, to **charge/radius** (e.g., "higher charge density → stronger attractions → higher mp")

Common Question

Explain the decrease in radius from K to K^+.
[2 marks]

Outline, in terms of their electronic structures, why the ionic radius of the phosphide ion is greater than that of the nitride ion.
[1 mark]

IB CHEMISTRY S2.2
THE COVALENT MODEL

Formation of covalent bonds

- **Covalent bond** = electrostatic attraction between a shared electron pair and two positively charged nuclei.
- **Octet rule**: many main-group atoms tend toward 8 valence electrons (duet for H).
- Some valid molecules have **<8 on the central atom** (e.g., BF_3) or an odd total (e.g., NO).
- **Lewis formulas** show **all valence e⁻** (bonding and lone pairs); use **dots/crosses/dashes**.

Common Mistake

Don't forget to add/subtract electrons for **charged ions** (e.g., +1 removes one electron).

Single/double/triple bonds: length vs strength

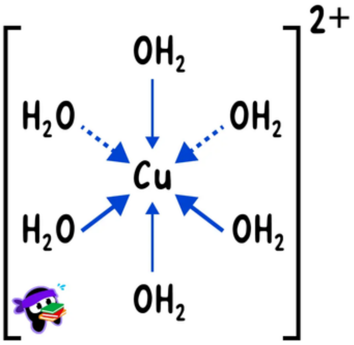
- **Trend** (same atom pair): more shared pairs → shorter and stronger bond.
- **Length**: single > double > triple
- **Strength**: triple > double > single
- **Reason**:
  - more **electron density** between nuclei increases attraction and draws nuclei closer;
  - higher **bond order** → higher bond **energy**.

Coordination bonds

- **Coordination bond**: both electrons in the shared pair originate from the same atom (the donor/Lewis base) and are **accepted** by an **electron-deficient** atom/ion (Lewis acid).
- Once formed, it's still a **covalent bond**.
- **Spotting in formulas/structures**: look for a lone-pair donor → to an electron-poor center (H^+, B, transition metal).
- Often drawn with an **arrow** (e.g., N→H).

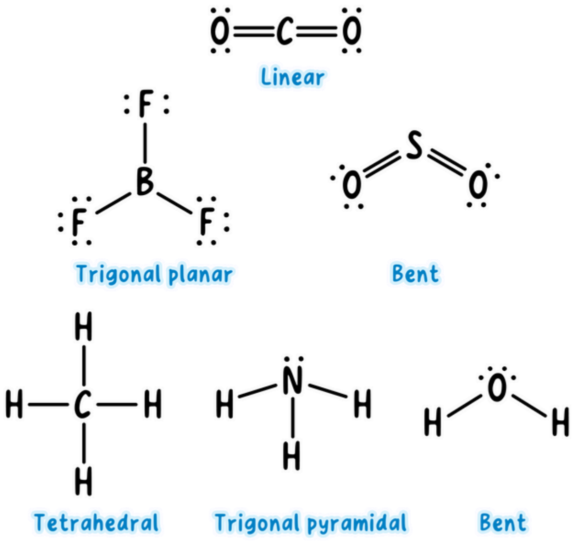
Exam Tip!

Keep in mind that coordination bonds often create **cations** (e.g., NH_4^+) without changing the octet on the donor.



VSPR theory

- **Rules**: **electron domains** (bonding pairs - single/multiple count as ONE domain - and lone pairs) repel; **geometry** minimizes repulsion.
- **Lone pairs** repel more than bonding pairs; **multiple** bonds exert slightly more repulsion than single.
- **Electron domain geometry → Molecular geometry (central atom) → Typical angles**:
  - **2** domains: linear →
    - **linear** (AX_2) → 180°
  - **3** domains: trigonal planar →
    - **trigonal planar** (AX_3, ~120°)
    - **bent** (AX_2E, ~120°→~117°)
  - **4** domains: tetrahedral →
    - **tetrahedral** (AX_4, 109.5°)
    - **trigonal pyramidal** (AX_3E, ~107°)
    - **bent** (AX_2E_2, ~104.5°)

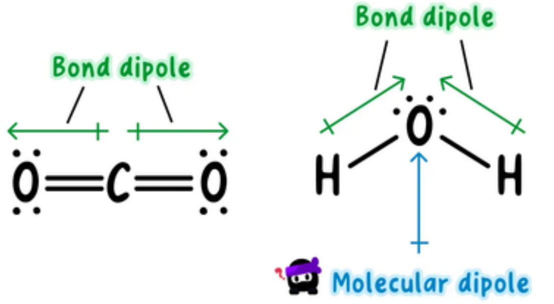


Bond polarity

- **Bond polarity** arises from ΔEN between atoms; higher ΔEN → more polar bond; extremes give ionic bonding.
- **Notation**: δ^+—δ^- (partial charges) and/or dipole vector arrow (→ toward δ^-).

Exam Tip!

Use data booklet EN values **qualitatively**: ΔEN ≲ ~0.4 (nonpolar), ~0.5–1.7 (polar covalent), larger tends ionic, but boundaries are fuzzy.



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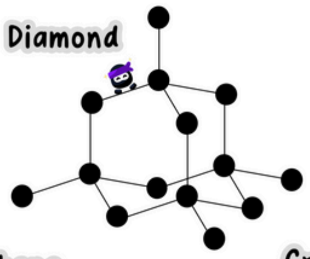
Molecular polarity

- **Rule**: a molecule/ion is polar if its bond dipoles do not cancel (net dipole ≠ 0). Symmetry often cancels dipoles.
- **Canonical outcomes** (≤4 domains):
    - **Linear** AX_2 (CO_2): nonpolar (equal & opposite).
    - **Trigonal planar** AX_3 (BF_3): nonpolar; AX_2E (SO_2): polar.
    - **Tetrahedral** AX_4 (CCl_4): nonpolar; AX_3E (NH_3): polar; AX_2E_2 (H_2O): polar.
    - $CHCl_3$ vs CCl_4: $CHCl_3$ is polar (vectors don't cancel due to H).

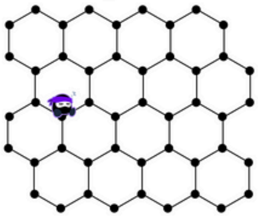
Exam Tip!

Draw vectors on each bond, then sum; mention symmetry (e.g., "tetrahedral but with one different substituent → net dipole")

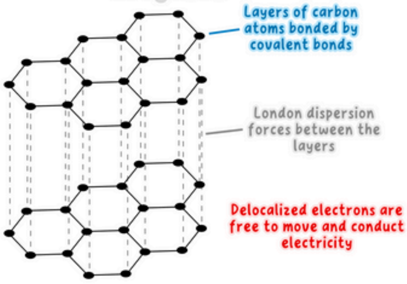
Diamond



Graphene



Graphite



Intermolecular forces (IMF)

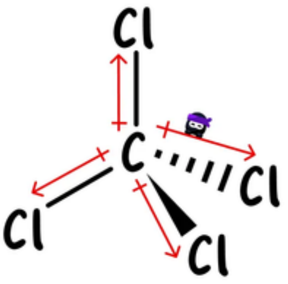
Use **"van der Waals forces"** as an inclusive term for:

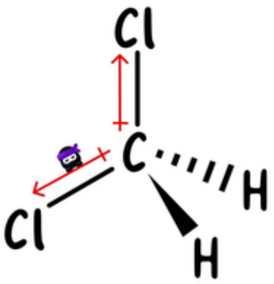
- **London (dispersion)**: present in all molecules/atoms; strength ↑ with molar mass & surface area (linear > branched).
- **Dipole-dipole**: between permanent dipoles (polar molecules).
- **Dipole-induced dipole**: polar molecule induces a dipole in a non-polar neighbor.
- **Hydrogen bonding** (H-bonding): a particularly strong dipole–dipole when H is covalently bound to N, O, or F and interacts with a lone pair on N/O/F of a neighboring molecule (X–H…Y).

Exam Tip!

To deduce IMFs quickly:

1. **Any N-H/O-H/F-H?** → H-bonding + dipole-dipole + London.
2. **Polar bonds arranged asymmetrically?** → dipole-dipole + London.
3. **Non-polar overall?** → London (and possibly dipole-induced dipole with polar neighbors).





Covalent networks

- **Diamond** (sp^3): each C tetrahedrally bonded to 4 C; giant 3D network → extremely hard, very high mp, electrical insulator, thermally conductive, insoluble.
- **Graphite** (sp^2): layers of hexagonal sheets; within a sheet strong σ + delocalized π; between sheets weak forces → layers slide (lubricant); conducts electricity/heat along planes; opaque, soft.
- **Graphene**: single graphite layer; exceptional electrical/thermal conductivity, very strong yet flexible, high surface area.
- **Fullerenes** (**C_{60}**): discrete molecular cages; simple molecular solid → lower mp, soluble in nonpolar solvents, semiconducting behavior.
- **Silicon** (sp^3 network) & **Silicon dioxide** (SiO_2): giant covalent lattices; high mp/hard; Si is a semiconductor (temperature-dependent conductivity); SiO_2 is an electrical insulator, hard, insoluble; tetrahedral Si–O–Si network.

Common Mistake

Don't say that "graphite is soft because bonds are weak". In reality, **in-plane** bonds are **strong**; softness is due to **weak interlayer forces**.

IMF strength order & properties

- **General strength** (similar molar mass): London < dipole–dipole < H-bonding (network covalent ≫ any IMF).
- **Volatility/boiling point**: stronger IMF → lower volatility, higher b.p. (branched isomers boil lower than straight chains).
- **Electrical conductivity**: simple molecular substances do not conduct (no mobile ions/electrons); acids/bases conduct in solution via ions.
- **Solubility** ("like dissolves like"): polar/H-bonding solutes dissolve in polar solvents (e.g., alcohols in water); non-polar solutes dissolve in non-polar solvents (e.g., hexane + iodine).

Common Question

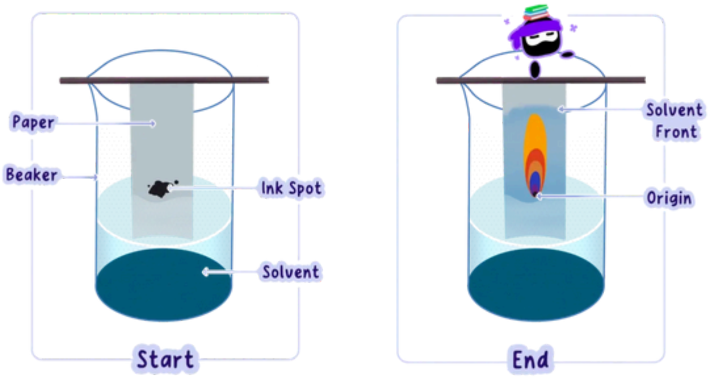
Which molecule is polar?

A. CO_2
B. BF_3
C. CCl_4
D. SO_2

Chromatography

- **Principle**: separation by different attractions (IMFs) to stationary phase vs mobile phase.
- **Interpretation** (silica/alumina = polar stationary phase, non-polar mobile):
  - more polar solute → lower Rf (sticks to plate)
  - less polar → higher Rf.

$$R_F = \frac{\text{distance traveled by solute}}{\text{distance traveled by solvent front}}$$

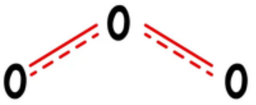


Resonance structures (HL)

- **Resonance structures**: two or more valid Lewis formulas differing only in position of π bonds and/or lone pairs (atom positions unchanged).
- The true structure is a delocalized **resonance hybrid**;
- **Bond order** often **fractional**
- All contributing forms linked by a **double-headed arrow** (↔).
- **Examples**: NO_3^-, NO_2^-, CO_3^{2-}, O_3, benzene.

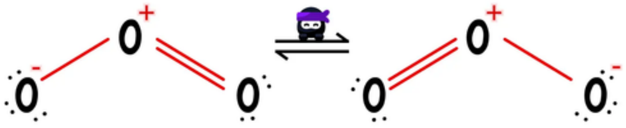
Benzene (HL)

- **Bond lengths**: all C–C bonds identical (~139 pm), intermediate between C-C and C=C.
- **Thermochemistry**: heat of hydrogenation << 3 × (cyclohexene); indicates resonance stabilization.
- **Chemical behavior**: prefers electrophilic substitution over addition (preserves aromatic system).
- **Spectroscopy/shape**: planar hexagon, ~120°; one ^1H NMR signal (all H equivalent).



Ozone hybrid

Benzene hybrid

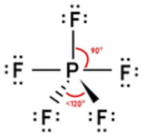


Ozone resonance structures

Benzene resonance structures

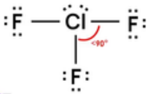
Expanded octets & VSEPR (HL)

- Period 3+ centers can **exceed an octet.**
- **Count electron domains** (bonds = 1 domain; lone pair = 1).
- **Five domains** → trigonal bipyramidal ED geometry (axial 180°, equatorial 120°, axial–equatorial 90°):
  - **5** bonding: PF_5 (TBP).
  - **4** bonding + 1 LP: SF_4 (see-saw, LP equatorial).
  - **3** bonding + 2 LP: ClF_3 (T-shaped, both LP equatorial).
  - **2** bonding + 3 LP: ICl_2^- (linear).
- **Six domains** → octahedral ED geometry (all 90°/180°):
  - **6** bonding: SF_6 (octahedral).
  - **5** bonding + 1 LP: BrF_5 (square pyramidal).
  - **4** bonding + 2 LP: XeF_4 (square planar; LP trans).

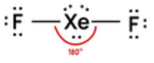


Trigonal bipyramid

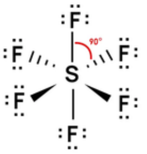
Seesaw



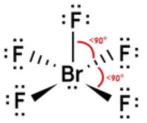
T-shape



Linear



Octahedral



Square pyramid

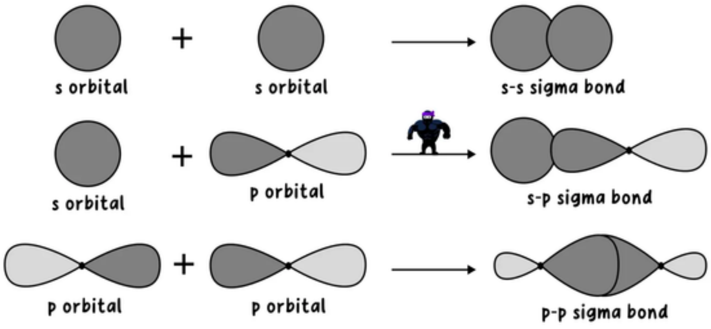
Square planar

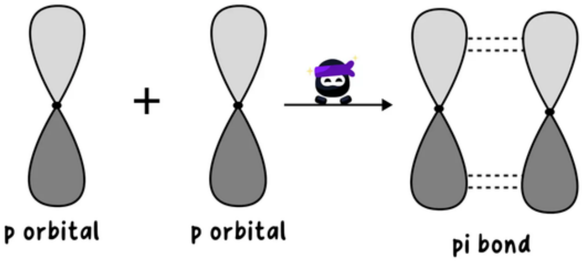
Formal charge (FC) (HL)

- **Formula**: FC = valence e^- – [nonbonding e^- + ½(bonding e^-)].
- **Preferred structure** tends to minimize charge, place negative FC on more electronegative atoms, and match the overall charge.

σ and π bonds (HL)

- **σ bond**: head-on overlap along internuclear axis (s–s, s–p, p–p, or hybrid–hybrid).
- **π bond**: side-on overlap of unhybridized p orbitals; electron density above/below bond axis.
- **Rule**: single bond = 1 σ; double = 1 σ + 1 π; triple = 1 σ + 2 π.

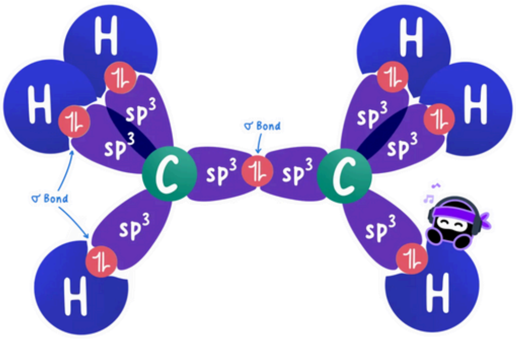




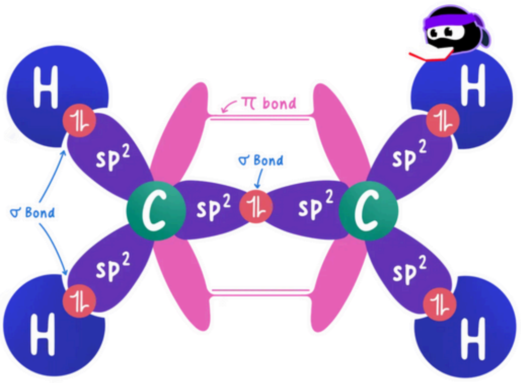
Hybridization (sp, sp^2, sp^3) (HL)

- Map **central-atom electron domains** to **hybridization & shape**:
  - **sp (2 domains)**: linear (180°) — e.g., $BeCl_2$, $HC{\equiv}CH$ (each C is sp; two π bonds from two unhybridized p's).
  - **sp^2 (3 domains)**: trigonal planar (120°) — e.g., BF_3, $H_2C{=}CH_2$ (each C sp^2 with one unhybridized p giving the π bond).
  - **sp^3 (4 domains)**: tetrahedral (109.5°) — e.g., CH_4 (AX_4), NH_3 (AX_3E), H_2O (AX_2E_2).
- **Resonance tip**: atoms adjacent to π systems often adopt sp^2 to allow p–p overlap (e.g., amide N is effectively sp^2, planar).
- **Predicting workflow**: Lewis → count domains (lone pairs included) → assign hybridization → predict geometry; or given geometry → infer hybridization.

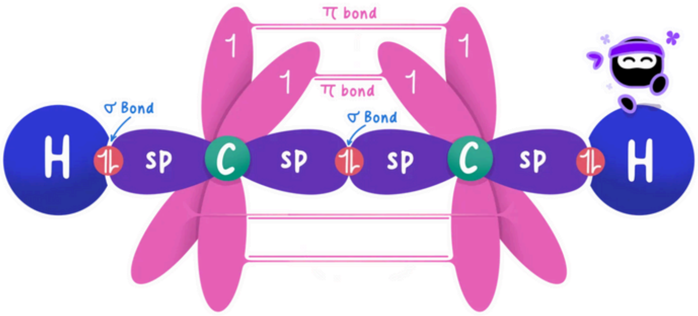
sp^3 hybridization



sp^2 hybridization



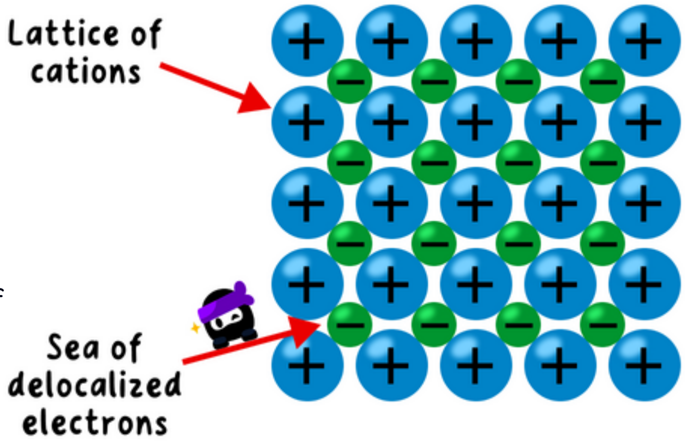
sp hybridization



IB CHEMISTRY S2.3
THE METALLIC MODEL

Metallic bonding & key properties

- **Metallic bond** = electrostatic attraction between a regular lattice of metal cations and a “sea” of delocalized valence electrons
- **Electrical conductivity**: delocalized electrons move under an applied potential → metals conduct in solid and liquid states
- **Thermal conductivity**: mobile electrons transfer kinetic energy; vibrations of the lattice(phonons) also carry heat → metals are good heat conductors
- **Malleability/ductility**: non-directional bonding; layers of ions can slide without breaking the overall attraction to the electron sea → metals can be hammered into sheets/drawn into wires
- Uses linked to properties:
  - **Copper** for electrical wiring (high conductivity, ductile);
  - **Aluminum** for heat sinks/foil (light, conductive, malleable);
  - **Gold** for contacts (corrosion-resistant, conductive);
  - **Alloys** for strength/hardness (controlled by disrupting lattice with different-sized atoms)



Exam Tip!

When explaining a **property**, name the micro-cause explicitly: “delocalized electrons” for conductivity, “non-directional attraction to electron sea” for malleability

Factors affecting metallic bodn strength & melting point trends

- **Bond strength increases** with **cation charge** (higher positive charge attracts electrons more strongly) and **decreases** with **larger ionic radius** (lower charge density)
- **Electron density**: higher number of delocalized electrons per atom strengthens bonding (e.g., Al has 3 vs Na has 1)
- **Down a group (s-block)**: ionic radius increases, charge unchanged → weaker bonding → melting points generally decrease
  - Group 1: Li > Na > K > Rb > Cs
  - Group 2: Be > Mg > Ca > Sr > Ba)
- **Across a period (s → p metals)**: In s/p blocks, increasing charge and electron contribution typically strengthen bonding → melting points tend to rise from Group 1 to Group 13 metal (e.g., Na < Mg < Al in Period 3), before dropping for post-metal p-block elements where metallic bonding weakens

Exam Tip!

Phrase **trend** answers as “greater charge and smaller radius → greater electrostatic attraction between cations and electron sea → higher m.p./hardness”

Transition elements: delocalized d-electrons, high melting point, conductivity (HL)

- **Transition metals** have **partially filled d subshells**; their d electrons contribute to the delocalized electron pool → higher electron density and strong metallic bonding
- **High melting points**: stronger cation–electron attraction from greater charge density and additional d-electron participation (e.g., Fe, Ni, W with very high m.p.)
- **Excellent electrical conductivity**: overlapping s and d bands provide many available states for electrons to move into under a potential difference
- **Mechanical properties**:
  - Generally **hard/tough** due to strong bonding
  - **Amenable to alloying** because d-electrons allow varied metallic radii and bonding types

IB CHEMISTRY S2.4 FROM MODELS TO MATERIALS

Bonding continuum & the bonding triangle

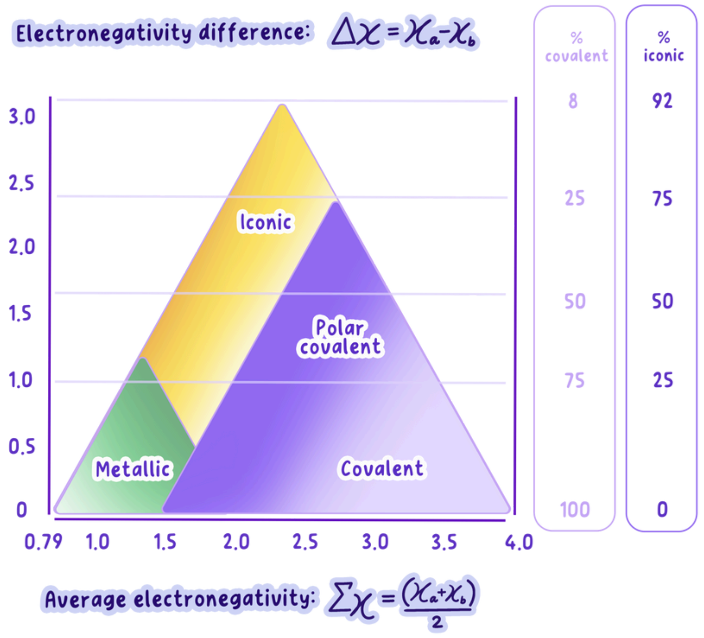
- Bonding lies on a **continuum** between ionic, covalent, and metallic; many solids have **mixed character**
- **Bonding triangle** (use data booklet):
  - corners="ideal" ionic/covalent/metallic
  - real materials plot inside based on mixed contributions
- Use **models → properties**:
  - **ionic** (lattice, strong electrostatics) → high mp, brittle, conducts when molten/aq;
  - **covalent** (network/molecular) →
    - network: very hard, very high mp, insulator
    - simple molecular: low mp/soft
  - **metallic** → conductive, malleable, variable mp

Position in the bonding triangle from electronegativity & predicting properties

- **Placement idea**:
  - **larger EN difference** (ΔEN) between two elements → more ionic character
  - **similar EN** → more covalent
  - presence of **metal atoms** contributing delocalized e⁻ → metallic character
- How to **place a binary compound quickly**:
  - identify metal/non-metal
  - estimate ΔEN from booklet
  - consider polarizability (size/charge density) to adjust toward covalent if ions are highly polarizing
- **Predict properties from position**:
  - more **ionic** → higher mp, brittle, soluble in polar solvents
  - more **covalent** (**molecular**) → low mp/volatile
  - more **covalent** (**network**) → very hard/insulating (semiconducting for some)
  - more **metallic** → electrical/thermal conductivity, malleability

Polymers

- **Polymers** = macromolecules from repeating monomers; properties arise from chain length, branching, polarity, cross-linking, crystallinity
- **Common plastic traits**: low density; soft to tough; electrical insulators
- **Microstructure**:
  - **more branching** → lower crystallinity → softer, clearer
  - **linear/regular chains** → higher crystallinity → stronger, opaque



Alloys

- **Alloys** = metal + other metals/non-metals;
- **Non-directional metallic bonding** remains, but lattice is disrupted by different-sized atoms → dislocation motion hindered
- **Property boosts**:
  - strength/hardness ↑ (dislocation pinning),
  - corrosion resistance ↑ (passivating oxides, Cr_2O_3 in stainless)
  - toughness/tailored conductivity/melting point changes

Common Mistake

Don't think that alloying "breaks" metallic bonding: delocalized electrons still bind the lattice, so we just impede layer motion

Common Question

Adding carbon atoms into the lattice of iron forms distortions that prevent the layers of iron atoms from sliding easily. Which property of the resulting alloy increases as a result?

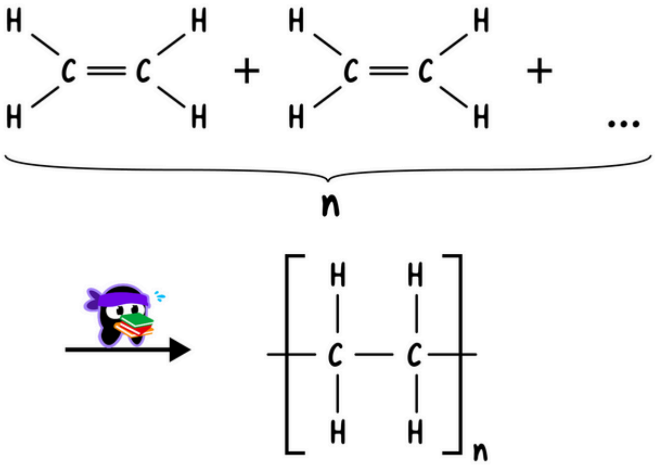
A. Electrical conductivity
B. Ductility
C. Hardness
D. Malleability

Addition polymers

- **Mechanism idea**:
  - **π bond** of each alkene **breaks**
  - **new σ bonds** link monomers into a chain
  - no small molecule eliminated
- **Repeating-unit notation**:
  - Draw the two-carbon backbone from the alkene with substituents, put in **brackets with "n" subscript**
  - **Side groups preserved** from monomer

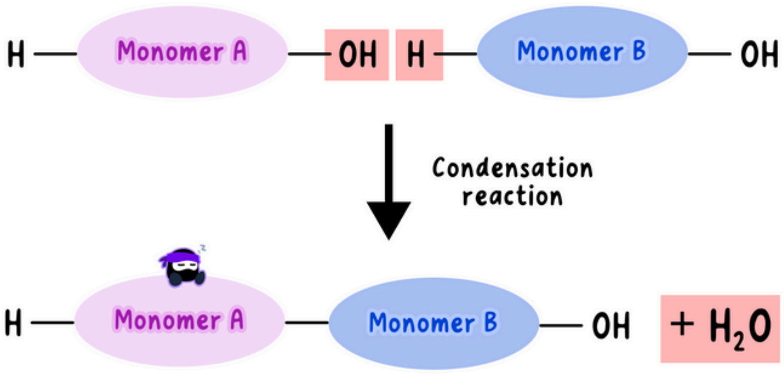
Exam Tip!

Start from the alkene, erase the double bond, extend two new single bonds left/right through brackets to indicate continuation.



Condensation polymers (HL)

- **Condensation polymerization**: functional groups on monomers react to form links while eliminating a small molecule (usually H_2O or HCl)
- **Polyamides** (diamine + diacid chloride/diacid):
  - amide (–CONH–) link
  - e.g., Nylon-6,6 from hexamethylenediamine and adipic acid → repeating unit $-NH-(CH_2)_6-NH-CO-(CH_2)_4-CO-$
- **Polyesters** (diol + diacid):
  - ester (–COO–) link
  - e.g., PET from ethane-1,2-diol and terephthalic acid → repeating unit $-O-CH_2-CH_2-O-CO-C_6H_4-CO-$
- **Repeating-unit notation**:
  - Draw between brackets the **linkage fragment** with end bonds extended;
  - Do not include the eliminated small molecule in the unit



Exam Tip!

- Circle the two atoms that join
- Write the linkage (–CONH– or –COO–) in the bracket
- Add "+ nH_2O" (or HCl) to balance if asked for an equation

IB CHEMISTRY S3.1 THE PERIODIC TABLE: CLASSIFICATION OF ELEMENTS

Periods, groups, blocks; metals/metalloids/non-metals

- **Periods** = horizontal rows; **groups** = vertical columns (1–18)
- **Blocks** by last-filled **sublevel**: s-block (G1–2+He), p-block (G13–18), d-block (transition), f-block (lanthanides/actinides)
- **Metals**: left/bottom of staircase
- **Non-metals**: right/top
- **Metalloids** on staircase (B,Si,Ge,As,Sb,Te,Po*)
- **General properties**:
  - **metals** (conductive,malleable,form cations);
  - **non-metals** (insulators,brittle as solids,form anions/share e^-);
  - **metalloids** = intermediate/semiconducting

Period number → outer level; group → valence

- **Period number** = n of outermost occupied main level; group for s/p elements → number of valence e^-
- **From position to configuration**:
  - Period 3, G14 (Si) → $[Ne]3s^2 3p^2$
  - Period 4, G2 (Ca) → $[Ar]4s^2$
  - Period 4, d-block (G8 Fe) → $[Ar]3d^6 4s^2$
- **From configuration to position**: count outer ns/np for s/p groups; last e-in d→transition group
- **Known families**: alkali metals(G1), halogens(G17), transition elements(d-block), noble gases(G18)

Note!

You can be assessed on deducing electron configuration of the element up to Z=36.

Periodicity: atomic/ionic radius, IE, EA, EN

- **Across a period** (→):
  - atomic radius ↓ (Zeff ↑ pulls e^- closer)
  - first IE ↑ generally
  - EA becomes more negative
  - EN ↑
  - metallic character ↓
- **Down a group** (↓):
  - atomic radius ↑ (more shells)
  - first IE ↓
  - EA less exothermic
  - EN ↓
  - metallic character ↑
- **Ionic radius**:
  - anions > parent atom (coulombic expansion)
  - cations < parent atom (loss of a shell/increased Zeff per e^-)

Exam Tip!

Link **property → cause** explicitly, e.g., “IE increases across due to smaller radius and greater Zeff”

Group trends

- **G1 reactivity increases down the group**; reaction with water:
  - $2M(s)+2H_2O(l) \rightarrow 2M^+(aq)+2OH^-(aq)+H_2(g)$
  - Li slow → K very vigorous
- **G17 (halogens) reactivity decreases down the group**: halogen higher in group oxidizes halide of lower one
  - Cl_2 displaces Br^-/I^-
  - Br_2 displaces I^-

Example:

$Cl_2(aq)+2Br^-(aq) \rightarrow 2Cl^-(aq)+Br_2(aq)$

$2Na(s)+2H_2O(l) \rightarrow 2NaOH(aq)+H_2(g)$

Continuum of metallic and non-metallic properties

- **Basic metal oxides** (Group 1 and 2):
  - If M = Group 1:
    - $MO+H_2O \rightarrow M^+(aq)+OH^-(aq)$
    - $M_2O+H_2O \rightarrow 2MOH$
  - If M = Group 2:
    - $MO+H_2O \rightarrow M(OH)_2$
- **Amphoteric oxides** (e.g., Al_2O_3,ZnO): react with acids and bases
- **Acidic non-metal oxides**:
  - $CO_2+H_2O \rightleftharpoons H_2CO_3(aq)$
  - $SO_2+H_2O \rightarrow H_2SO_3$
  - $SO_3+H_2O \rightarrow H_2SO_4$;

    → non-metal oxides form acids
- **Environmental**: acid rain from $SO_2/SO_3/NO_x$; ocean acidification: $CO_2(g) \rightleftharpoons CO_2(aq)+H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$(→ $CaCO_3$ dissolution)

Oxidation states

- **Oxidation state (OS)** charge an atom would have if bonds were ionic
- **Key rules**:
  - elemental forms OS = 0 (e.g., Na(s), $O_2(g)$, $P_4(s)$);
  - monoatomic ion OS = its charge;
  - sum of OS in neutral compound = 0(in ion=ion charge)
- **Some of the common OS**:
  - Group1: +1;
  - Group2: +2;
  - F: −1;
  - H: +1 (except metal hydrides:−1);
  - O: −2 (except peroxides:−1; with F positive)

Common Question

Which property of elements decreases down a group but increases across a period?

A. Metallic character

B. Electronegativity

C. Atomic radius

D. Atomic mass

First-IE discontinuities as evidence for sublevels (HL)

- **General trend across a period**: IE_1↑ as Zeff↑ and atomic radius↓, but two predictable dips each short period arise from orbital energy and e^-–e^- repulsion
- **Be → B dip**: removed e^- is from higher-energy 2p (B) vs 2s (Be) → less energy required → $IE_1(B)<IE_1(Be)$
- **N → O dip**: in O the fourth 2p electron pairs with an existing one in a 2p orbital → intra-orbital repulsion raises its energy → easier to remove → $IE_1(O)<IE_1(N)$
- **Period 3 parallels**: Mg → Al and P → S show the same reasons (3s → 3p energy step; first pairing in 3p)
- Therefore, the pattern demands **sublevels** (s and p) of **different energies** and the concept of e^- pairing within degenerate orbitals

Common Mistake

Don't explain the **discontinuities** with "half-filled is especially stable" alone; IB expects an **energy-of-removed-electron** argument: higher sublevel energy and pairing repulsion.

Transition elements (HL)

- **Transition elements** are d-block elements Sc → Zn with an **incomplete d sublevel** in the element or common ions
- **Variable oxidation states**: successive removal of similar-energy 4s/3d electrons allows multiple OS (e.g., Fe^{2+}/Fe^{3+}, Mn(II)–Mn(VII))
- **High melting points**: strong metallic bonding with d-electron participation → high charge density, many mobile electrons
- **Magnetic properties** (knowledge of types is not assessed):
  - **unpaired d electrons** → paramagnetism(common)
  - **paired** → diamagnetism
- **Catalytic properties**: variable OS and partially filled d orbitals facilitate **adsorption**, **bond weakening**, and **electron transfer** at surfaces/complexes
- **Coloured compounds**: ligand fields **split d-levels**; d–d or charge-transfer absorptions give colors
- **Complex ion formation**: metal centers accept **lone pairs** from ligands to form coordinate bonds (e.g., $[Cu(H_2O)_6]^{2+}$)

Exam Tip!

When asked "why **colored**/why **catalytic**," name "split d-sublevels/variable OS & adsorption" explicitly.

Variable oxidation states in transition elements (HL)

- **Reason for variable oxidation states in transition elements**: successive IEs for removing 4s and then 3d electrons are close in value → several oxidation states are accessible depending on conditions/ligands

Exam Tip!

To deduce electron configurations of ions, remove 4s before 3d; start from neutral atom (remember Cr, Cu exceptions) then subtract electrons.

Example:

To deduce Fe^{3+} from Fe $[Ar]3d^6 4s^2$, remove $4s^2$ then one 3d → $[Ar]3d^5$.

Coloured transition element complexes (HL)

- **Cause of color**:
  - Ligand field **splits the five d orbital**
  - Absorbing a photon **promotes an electron** between split d levels
  - The **absorbed** band's color is **complementary** to the **observed** solution color
- **Complementary pairs** (use a simple wheel):
  - red ↔ green
  - orange ↔ blue
  - yellow ↔ violet
  - blue-green ↔ red-purple
- **Relating to wavelength λ and frequency f**:
  - $c=\lambda f$ and $E=hf$
  - shorter λ → higher f → higher E
  - **larger d-level splitting** → absorption shifts toward higher energy (blue/violet)

Example:

- $[Cu(H_2O)_6]^{2+}$ appears blue → absorbs orange (~600 nm)
- $[Ni(H_2O)_6]^{2+}$ appears green→absorbs red (~650–700 nm)
- $[Ti(H_2O)_6]^{3+}$ appears purple→absorbs yellow-green (~560 nm)

IB CHEMISTRY S3.2 FUNCTIONAL GROUPS: CLASSIFICATION OF ORGANIC COMPOUNDS

Representation of organic molecules

- **Empirical formula**: simplest whole-number ratio (e.g., CH_2O).
- **Molecular formula**: actual atom counts (e.g., $C_6H_{12}O_6$).
- **Structural formulas**:
  - **Full** (show all bonds); condensed (e.g., CH_3CH_2OH).
  - **Skeletal** (zig-zag lines for C–C; H on C omitted; heteroatoms & H on heteroatoms shown).
  - **Stereochemical** (when needed): wedge (out), dash (back) around a stereocenter.
- **Interconverting**: count carbons & heteroatoms → place on a skeletal backbone → expand to condensed/full as required.
- **3D models**: build mental/wedge-dash or kit/virtual for tetrahedral (sp^3), trigonal planar (sp^2), linear (sp) centers.

Homologous series

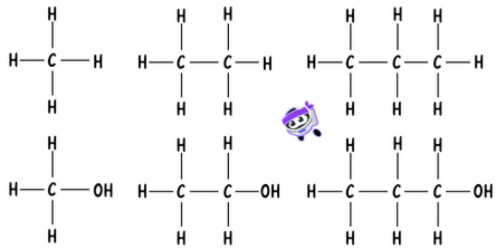
- **Alkanes** C_nH_{2n+2}
- **Alkenes** C_nH_{2n}
- **Alkynes** C_nH_{2n-2}
- **Halogenoalkanes** $C_nH_{2n+1}X$
- **Alcohols** $C_nH_{2n+2}O$ (R-OH)
- **Aldehydes** $C_nH_{2n}O$ (R-CHO)
- **Ketones** $C_nH_{2n}O$ (R-CO-R′)
- **Carboxylic acids** $C_nH_{2n}O_2$ (R-COOH)
- **Ethers** $C_nH_{2n+2}O$ (R-O-R′)
- **Amines** $C_nH_{2n+3}N$ (primary $R\text{-}NH_2$ baseline)
- **Amides** $C_nH_{2n}O$ N ($R\text{-}CONH_2$ baseline)
- **Esters** $C_nH_{2n}O_2$ (R-COOR′)

Functional groups

- **Halogeno** (R–X): X = F, Cl, Br, I.
- **Hydroxyl** (–OH): alcohols.
- **Carbonyl** (C=O): aldehyde (–CHO, end of chain), ketone (–CO–, within chain).
- **Carboxyl** (–COOH): carboxylic acids.
- **Alkoxy** (–OR): ethers R–O–R′.
- **Amino** ($-NH_2$ / –NHR / $-NR_2$): amines.
- **Amido** ($-CONH_2$/–CONHR/$-CONR_2$): amides.
- **Ester** (–COOR′).
- **Phenyl** ($-C_6H_5$): benzene ring as substituent.
- **Saturated vs unsaturated**: only C–C single bonds vs having C=C/C≡C.

Common Mistake

Don't mix **aldehyde** (–CHO) with **alcohol** ($-CH_2OH$): better check for C=O vs O–H next to carbon.



Physical trends in homologous series

- **Boiling/melting points:** increase with chain length (↑ London dispersion: greater surface area & polarizability).
- **Branching:** lowers b.p. (smaller surface area contact).
- **Hydrogen bonding:** raises b.p. (alcohols > alkanes of similar Mr).
- **Solubility**: in water decreases with longer carbon chains (hydrophobic effect) for polar groups; hydrocarbons are insoluble.

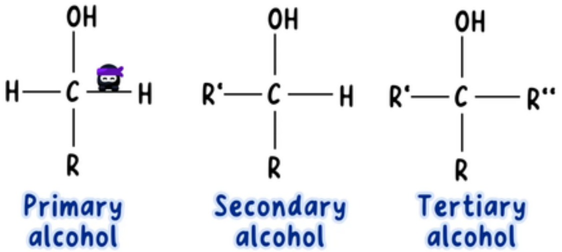
IUPAC nomenclature

Workflow:

- **Longest parent chain** containing the highest-priority functional group / C=C if present (≤ C_6).
- **Number** to give the **lowest locants** (position of a functional group or substituent) to the principal functional group and unsaturation.
- **Name substituents** (methyl, ethyl, fluoro/chloro/bromo/iodo, hydroxy, oxo for =O in aldehyde/ketone contexts).
- **Assemble**: prefix (substituents, alphabetized; di-, tri- ignored for ordering) + parent + infix for unsaturation (-ene) + suffix for main functional group (-ol, -al, -one, -oic acid). Include locants.

Structural isomerism

- **Chain (skeletal) isomers**: different carbon skeletons (butane vs 2-methylpropane).
- **Position isomers**: same functional group, different position (propan-1-ol vs propan-2-ol; bromo on C-1 vs C-2).
- **Functional-group isomers**: different functional groups (C_3H_6O: propanal vs propanone).
- **Primary (1°), secondary (2°), tertiary (3°) classification** (where applicable):
  - **Alcohols & halogenoalkanes**: by the number of carbons attached to the functional carbon (–C–OH or –C–X).
  - **Amines**: by the number of carbon groups on N (RNH_2 1°, R_2NH 2°, R_3N 3°).

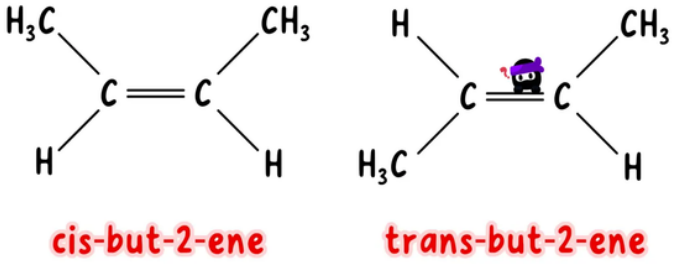


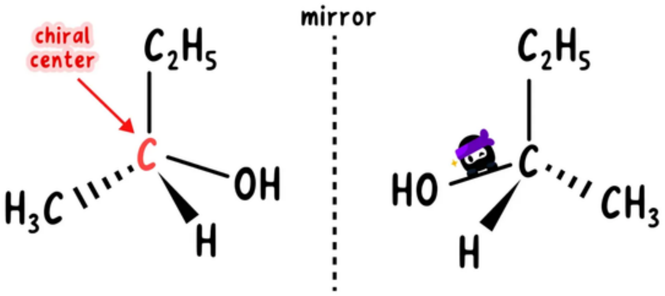
Stereoisomerism: cis-trans and chirality (HL)

- **Stereoisomers**: same connectivity; different 3D arrangement.
- **Cis-trans (geometric) isomerism** requires restricted rotation and each end of the unit having two different substituents:
- **Alkenes**: C=C with each sp^2 C bearing two different groups → cis (same side) vs trans (opposite) across the double bond.
- **Cycloalkanes** (C_3, C_4): ring prevents rotation; cis = substituents on same face; trans = opposite faces.
- **Chiral carbon (stereocenter)**: sp^3 C bonded to four different groups → two enantiomers (non-superimposable mirror images).
- **Draw with wedge-dash** (solid wedge = out of plane; dashed wedge = behind).
- **Optical activity**: enantiomers rotate plane-polarized light in equal and opposite directions; racemic mixtures are optically inactive.
- **Recognition**: superimpose models (real/virtual). If one cannot be rotated to match the other → enantiomers.

Exam Tip!

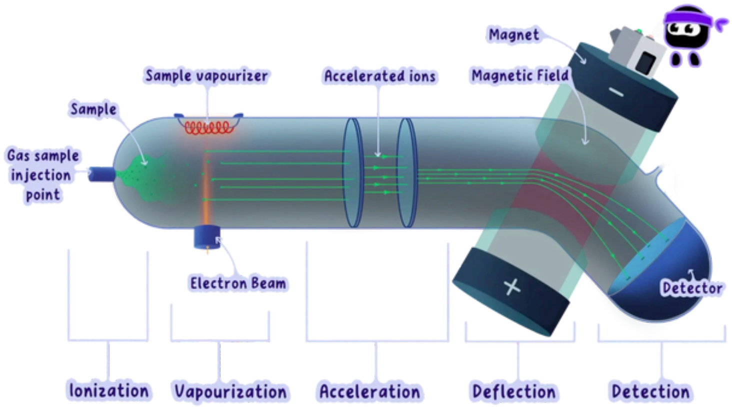
For C=C, quickly check "two different groups on each carbon". If not, no cis-trans possible.





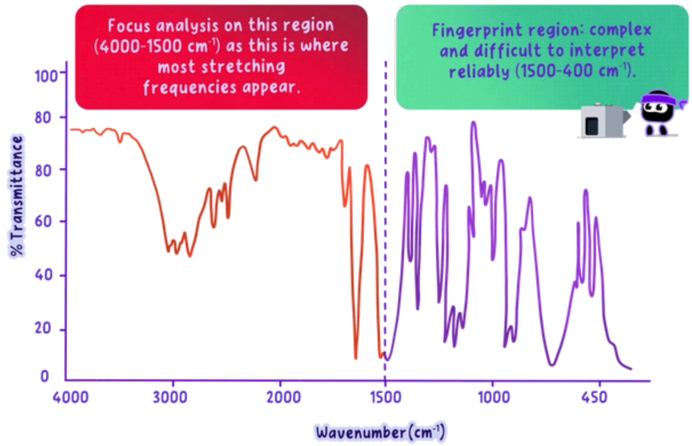
Mass spectrometry (MS) (HL)

- **Molecular ion ($M^{+\cdot}$)**: the highest-m/z isotopic cluster corresponding to the intact molecule (often weak/absent for easy-fragmenting compounds).
- **Base peak**: tallest peak (100% intensity) = most stable/abundant fragment.
- **Data booklet** gives **key m/z**
- **Strategy**: use M^+ for formula clue, isotope pattern for heteroatoms, diagnostic fragments for functional groups.



IR spectroscopy (HL)

- **Functional-group region (≈ 4000–1500 cm^{-1})**:
    - **O–H (alcohol)**: broad 3200–3600; O–H (acid): very broad 2500–3300 + C=O.
    - **N–H**: 3300–3500, sharper (1° amine: two bands; 2°: one).
    - **C=O**: strong ~1700 (aldehyde/ketone ~1710–1725; acid ~1710; ester ~1735–1750; amide ~1650–1700).
    - **C≡N/C≡C**: ~2200–2260 (sharp; C≡N stronger).
    - **C–O**: 1050–1300 (alcohol/ether/ester).
    - Other important regions are given in the data booklet
- **Fingerprint region (≤1500 cm^{-1})**: complex pattern; use to confirm identity with a known spectrum.
- **Greenhouse gases**: absorb IR via vibrations that change dipole moment (e.g., CO_2 asymmetric stretch, H_2O bends/stretch) → trap outgoing IR.



^{1}H NMR: *signals, chemical shift, integration* (HL)

- **Number of signals**: count distinct H environments (symmetry can merge sites).
- **Chemical shift (δ, ppm)**: environment clues (approx.):
- 0.8–1.2 alkyl CH_3; 1.2–1.7 CH_2; ~2–3 next to π/heteroatom (allylic, benzylic, α-to-C=O).
- ~3–4.5 next to O/N/halogen; ~4.5–6.5 vinylic (C=C–H); ~6.5–8 aromatic; ~9–10 aldehyde; ~10–12 acid (broad).
- **Integration**: area under a signal ∝ relative number of H in that environment (gives ratios).

^{1}H NMR splitting (n+1 rule) (HL)

- **Spin-spin coupling**: non-equivalent neighboring protons (usually within 3 bonds) split signals; n neighboring H → n+1 peaks.
  - **Singlet** (s): no neighbors (n=0).
  - **Doublet** (d): one neighbor (n=1).
  - **Triplet** (t): two neighbors (n=2).
  - **Quartet** (q): three neighbors (n=3).
- **Classic ethyl pattern**: $-CH_3$ (t) integrating 3H coupled to $-CH_2$; $-CH_2$ (q) integrating 2H coupled to $-CH_3$.

Combining data to solve structures (HL)

Workflow:

- **MS**: read M^+ (Mr) and isotopes → presence of halogens; note key fragments.
- **IR**: identify functional groups (C=O? O–H? N–H? C≡N?).
- **^{1}H NMR**: number of signals (symmetry), δ ranges (environment), integration (H count), splitting (connectivity).
- **Assemble pieces** to fit the formula/DBE (unsaturation count if needed), cross-check all spectra.

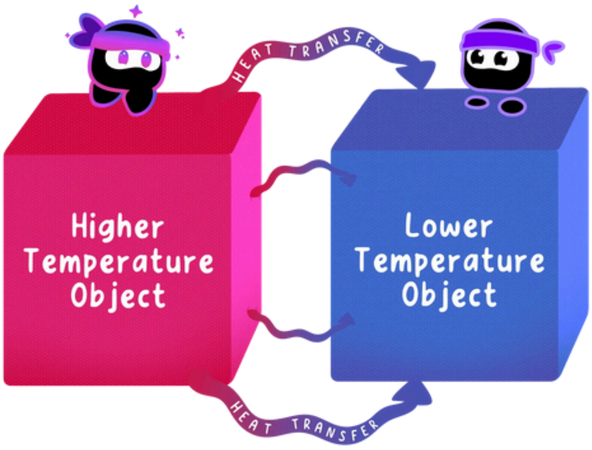
IB CHEMISTRY R1.1 MEASURING ENTHALPY CHANGES

Energy transfer; heat vs temperature

- **Heat** (q): **energy transfer** due to a **temperature difference**; path-dependent; flows system ⇄ surroundings
- **Temperature** (T): measure of **average kinetic energy** of particles; state property; determines direction of heat flow
- **System vs surroundings**
  - **System**: the chemicals that react;
  - **Surroundings**: everything else(often the solution, calorimeter, air)
- **Conservation of energy**: q(system) +q(surroundings) = 0

Note!

Unlike heat, **temperature** is a **state function**, meaning it depends only on the **current state** of the system, not how the system reached that state.



Endothermic vs exothermic; temperature changes

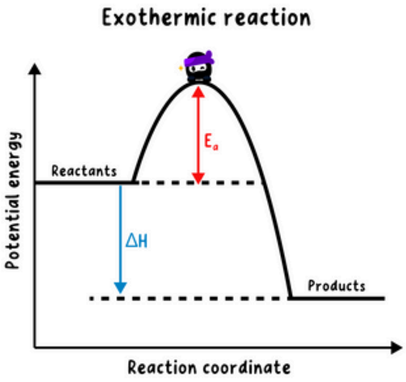
- **Endothermic**
  - System **absorbs** heat: **$\Delta H>0$**
  - Temperature of surroundings **decreases**
- **Exothermic**
  - System **releases** heat: **$\Delta H<0$**
  - Temperature of surroundings **increases**
- **Thermometer view**: In a simple solution calorimetry, thermometer reads surroundings' temperature change (≈solution)

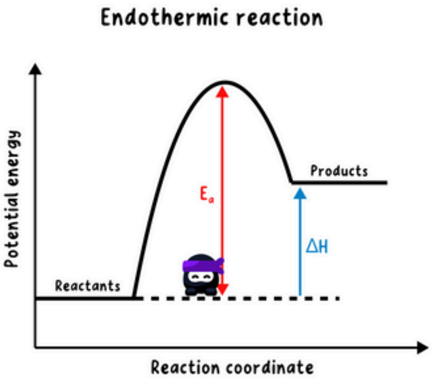
Common Question

Which describes an endothermic reaction?
A. Heat flows from the system to the surroundings, and the enthalpy of reactants is greater than that of products.
B. Heat flows from the surroundings to the system, and the enthalpy of reactants is greater than that of products.
C. Heat flows from the surroundings to the system, and the enthalpy of products is greater than that of reactants.
D. Heat flows from the system to the surroundings, and the enthalpy of products is greater than that of reactants.

Energy profiles; stability vs enthalpy

- **Energy profile axes**
  - x: reaction coordinate(progress)
  - y: potential energy
- **Exothermic** profile;
  - Products **lower** than reactants
  - $\Delta H=E(products)-E(reactants)<0$
- **Endothermic** profile;
  - Products **higher** than reactants
  - $\Delta H=E(products)-E(reactants)>0$
- **Relative stability**
  - **Lower** potential energy = ***greater*** thermodynamic ***stability***
  - **Exothermic** reactions form **more stable** products than reactants





Common Mistake

Don't forget to label **axes** and **ΔH arrow**: direction matters!

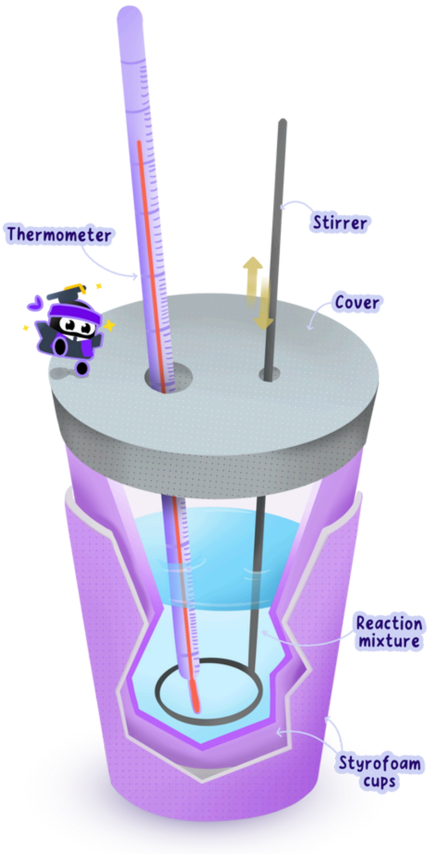
Standard enthalpy change ΔH⦵; calorimetry

- **ΔH⦵**: **enthalpy change** at constant pressure under standard conditions (usually 1.00×10^5Pa, 298K) and standard states
- **Calorimetry equations**
  - $Q = mc\Delta T$ (heat gained/lost by surroundings solution)
  - $\Delta H(\text{reaction}) = -Q(\text{surroundings}) / n(\text{imiting})$
- **Units & constants**
  - $c(\text{water}) \approx 4.18$ J g^{-1} K^{-1};
  - use m in grams to match units
  - report ΔH in kJ mol^{-1}
- **Sign convention**
  - If solution warms ($\Delta T>0$), $Q(\text{surroundings})>0$ → reaction released heat → $\Delta H<0$
  - If solution cools, $\Delta H>0$

Note!

Sources of error in calorimetry:

- **Heat loss to surroundings**: heat escapes from the calorimeter → the temperature change measured is less than the actual one
- **Incomplete combustion**: fuel may not burn completely in combustion reactions → less heat being released than expected
- **Specific heat capacity assumption**: for the solution, it is assumed that its specific heat capacity is the same as of pure water



IB CHEMISTRY R1.2 ENERGY CYCLES IN REACTIONS

Bond breaking/forming; average bond enthalpies

- **Bond-breaking** absorbs energy(+), **bond-forming** releases energy(−):
  - **overall** ΔH ≈ ΣE(bonds broken)−ΣE(bonds formed)
- **Average bond enthalpy** = mean energy to break a bond in the gas phase averaged over many compounds → context-dependent values **may differ from experimental ΔH** for a specific reaction
- **Procedure** (using average bond enthalpies):
  - Write balanced equation
  - Draw or list all bonds broken/formed (gas phase)
  - Sum energies
  - ΔH ≈ broken−formed (sign tells if endothermic or exothermic)

ΔHc° and ΔHf° definitions & use (HL)

- **Standard enthalpy of combustion, ΔHc°**: enthalpy change when 1 mol of substance burns completely in O_2 under standard conditions (typically negative)
- **Standard enthalpy of formation, ΔHf°**: enthalpy change when 1 mol of compound forms from its elements in their standard states (elements: ΔHf°=0)
- **Use**: build cycles or apply summation formulas to solve for unknown ΔH

Note!

Always check **states** (g, l, s, aq) match the table you're using, because ΔH values are **state-dependent.**

Hess's law and multistep routes

- **Hess's law**: ΔH depends only on initial and final states; enthalpy changes are path-independent
- **Method**: manipulate given equations (**flip** → change sign; **multiply** → scale ΔH) so they sum to target equation; then sum ΔH values

Example:

- Target: $C(graphite)+\frac{1}{2}O_2 \rightarrow CO(g)$
- Given:
  - (1) $C+O_2 \rightarrow CO_2$ ΔH_1
  - (2) $CO+\frac{1}{2}O_2 \rightarrow CO_2$ ΔH_2
- Reverse (2): $CO_2 \rightarrow CO+\frac{1}{2}O_2$ ($\Delta H=-\Delta H_2$);
- Add to (1): $C+O_2+CO_2 \rightarrow CO_2+CO+\frac{1}{2}O_2 \rightarrow C+\frac{1}{2}O_2 \rightarrow CO$
- $\Delta H=\Delta H_1-\Delta H_2$

Calculating ΔH° with ΔHf° or ΔHc° data (HL)

- From **formations** and **combustion** values:

$$\Delta H^\ominus = \Sigma\left(\Delta H_f^\ominus \text{ products }\right) - \Sigma\left(\Delta H_f^\ominus \text{ reactants }\right)$$
$$\Delta H^\ominus = \Sigma\left(\Delta H_c^\ominus \text{ reactants }\right) - \Sigma\left(\Delta H_c^\ominus \text{ products }\right)$$

Common Mistake

Don't mix ΔHf° and ΔHc° data in the **same summation**: stick to one type per calculation unless you're explicitly constructing a Hess cycle.

Born-Haber cycles (HL)

- **Purpose**: apply Hess's law to lattice formation; link macroscopic ΔHf° to atomic-scale steps

Typical steps for MX(s) from elements (sign conventions in brackets):

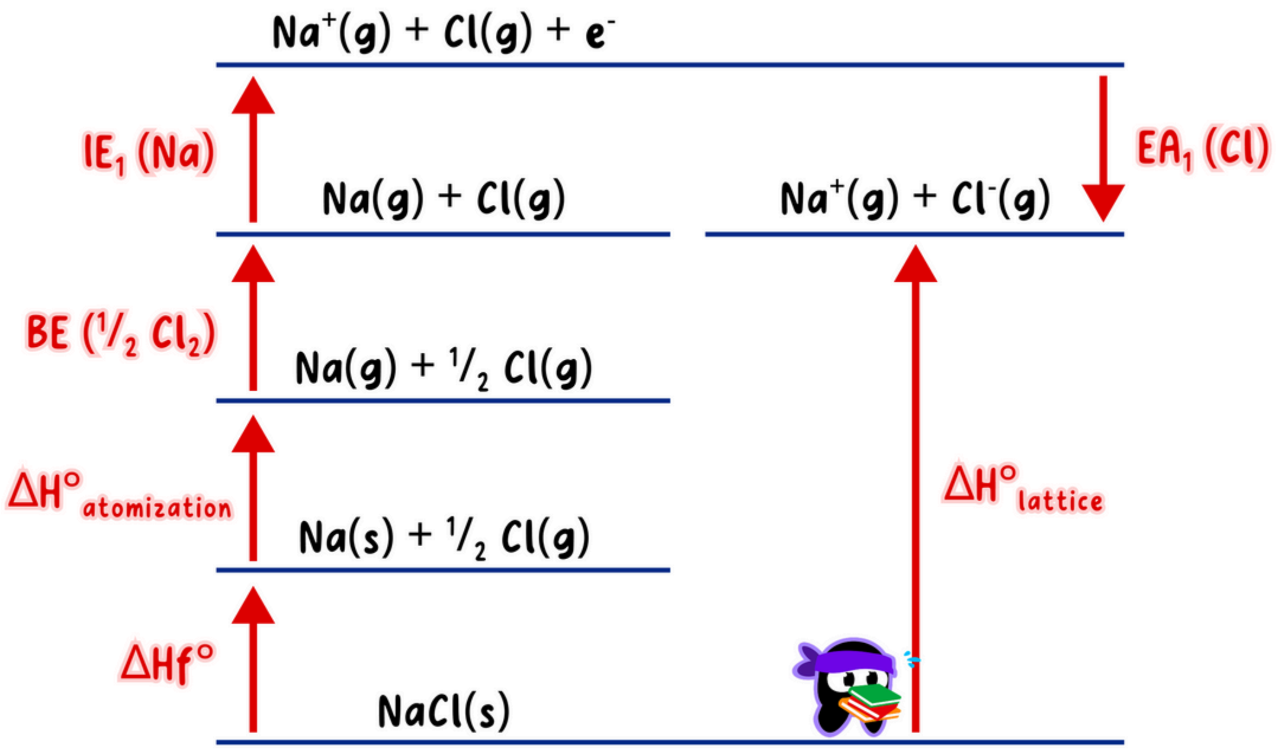
- **Atomization of metal** M(s) → M(g) [endothermic, +]
- **Atomization/dissociation of non-metal** $\frac{1}{2}X_2(g) \rightarrow X(g)$ (or bond enthalpy/2) [endothermic, +]
- **Ionization of metal** $M(g) \rightarrow M^+(g)+e^-$ (IE_1; add IE_2 for M^{2+} etc.) [endothermic, +]
- **Electron affinity of non-metal** $X(g)+e^- \rightarrow X^-(g)$ (EA_1 often −; EA_2 for O,S is +) [usually exothermic (EA_1 −)]
- **Lattice enthalpy**: $M^+(g)+X^-(g) \rightarrow MX(s)$ [exothermic, very −] (by IB convention often "formation from gaseous ions")
- **Cycle sums to standard formation**: $M(s)+\frac{1}{2}X_2(g) \rightarrow MX(s)$ with ΔHf° (usually −)

Note!

During IB examination, you will not be asked to construct a complete Born-Haber cycle, but you can be assessed on determining specific enthalpy values involved in the cycle.

Example:

Below you can see the Born-Haber cycle for **sodium chloride**.



IB CHEMISTRY R1.3 ENERGY FROM FUELS

Combustion of metals, non-metals, hydrocarbons & alcohols

- **Combustion** =
  - redox with O_2 producing oxides;
  - complete combustion of organics yields CO_2 and H_2O;
  - balance C→H→O last
- **Metals**:
  - $4Fe(s)+3O_2(g) \rightarrow 2Fe_2O_3(s)$
  - $2Mg(s)+O_2(g) \rightarrow 2MgO(s)$
- **Non-metals**:
  - $C(s)+O_2(g) \rightarrow CO_2(g)$
  - $S(s)+O_2(g) \rightarrow SO_2(g)$
  - $P_4(s)+5O_2(g) \rightarrow P_4O_{10}(s)$
- **Hydrocarbons** (general): $CxHy+O_2 \rightarrow xCO_2+(y/2)H_2O$
- **Alcohols** (general): $CxHyO+O_2 \rightarrow xCO_2+((y/2))H_2O$

Incomplete combustion → CO and C

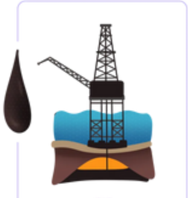
- **Incomplete combustion** occurs under O_2 limitation/poor mixing/low T; products include CO(g) and C(s) with H_2O
- **Hydrocarbons** examples:
  - **CO-rich**:
    - $2CH_4+3O_2 \rightarrow 2CO+4H_2O$
    - $2C_3H_8+7O_2 \rightarrow 6CO+8H_2O$
  - **Soot-rich**: (not unique; many stoichiometries possible)
    - $CH_4+O_2 \rightarrow C+2H_2O$
    - $C_3H_8+2O_2 \rightarrow 3C+4H_2O$
- **Alcohols** examples:
  - CO-rich: $2C_2H_5OH+3O_2 \rightarrow 4CO+6H_2O$
  - Soot-rich: $C_2H_5OH+O_2 \rightarrow 2C+3H_2O$

Biofuels, photosynthesis, renewable vs non-renewable

- **Biofuels** include bioethanol, biodiesel (FAME), biogas; “short-cycle” carbon from recent photosynthesis
- **Photosynthesis** (net): $6CO_2+6H_2O \rightarrow C_6H_{12}O_6+6O_2$ (sunlight, chlorophyll)
- **Renewable vs non-renewable**:
  - Renewable **replenishes** on human timescales (biofuels, solar, wind)
  - Non-renewable **does not** (coal, oil, gas)
- **Advantages of biofuels**: potentially lower net CO_2 over lifecycle, domestic production, biodegradable; some lower PM/CO with proper engines
- **Disadvantages**: land use/deforestation, food vs fuel, variable net GHG when considering fertilizers, processing, transport; lower energy density (ethanol vs gasoline), cold-start issues

Fossil fuels

- **Fuels & idealized CO_2 per gram** (complete combustion):
  - CH_4: 16 g → 44 g CO_2 → 2.75 g CO_2 g^{-1} (high H/C → least CO_2 per energy among hydrocarbons)
  - C_8H_{18}: 114 g → 8×44 = 352 g CO_2 → 3.09 g g^{-1}
  - Carbon(coal model): 12 g → 44 g CO_2 → 3.67 g g^{-1} (highest CO_2 g^{-1})
  - Ethanol C_2H_5OH: 46 g → 88 g CO_2 → 1.91 g g^{-1};
  - Methanol CH_3OH: 32 g→44 g CO_2 → 1.38 g g^{-1}
  - H_2: 0 g CO_2 at point of use
- **Incomplete combustion tendency**:
  - Higher with poor O_2 supply, low flame temperature, large droplets/poor atomization, soot-forming aromatics
- **Energy released per mass** (qualitative, typical):
  - H_2 > natural gas (CH_4) > gasoline range (C_5–C_{12}) > diesel/heavier fractions > coal
  - alcohols generally lower than comparable hydrocarbons (O already present reduces ΔH_c)
- **Fossil fuels**:
  - **Natural gas**: low CO_2 per kWh, clean burn, but CH_4 leakage is a potent GHG
  - **Oil/gasoline**: high energy density, infrastructure mature; higher CO_2 per kWh and pollutants without controls
  - **Coal**: abundant, cheap; highest CO_2 per kWh, SO_2/NO_x/PM unless scrubbed
- **Link to greenhouse effect**: CO_2 absorbs IR (vibrational modes) → enhanced greenhouse effect; rising atmospheric CO_2 correlates with increased radiative forcing and global warming

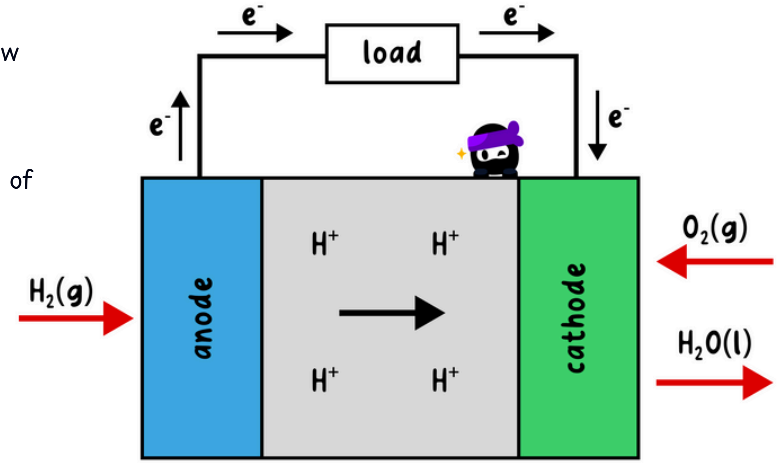

Oil


Coal


Gas

Fuel cells (H_2 and methanol)

- **Fuel cells** convert chemical energy directly to electrical via redox at electrodes; continuous fuel/oxidant supply; acidic media equations below
- **Hydrogen fuel cell** (acidic, PEM):
  - Anode(oxidation): $H_2 \rightarrow 2H^+ + 2e^-$
  - Cathode(reduction): $\frac{1}{2}O_2 + 2H^+ + 2e^- \rightarrow H_2O$
  - Overall: $H_2 + \frac{1}{2}O_2 \rightarrow H_2O(l)$ (no CO_2 at point of use)
- **Methanol fuel cell** (acidic, DMFC):
  - Anode: $CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$
  - Cathode: $(3/2)O_2 + 6H^+ + 6e^- \rightarrow 3H_2O$
  - Overall: $CH_3OH + (3/2)O_2 \rightarrow CO_2 + 2H_2O$
- **Properties/uses**:
  - High efficiency at part load, quiet, modular
  - H_2 FCs suited for vehicles/stationary backup
  - DMFCs for portable/low-power applications



Hydrogen fuel cell

Note!

The concept of redox reactions, which is also relevant to this subtopic, is studied in more detail in **R3.2 Electron transfer reactions.**

IB CHEMISTRY R1.4 ENTROPY AND SPONTANEITY (HL ONLY)

Entropy S, predicting ΔS, and using $S^\ominus$ tables (HL)

- Entropy idea: microstate count/dispersal of energy and matter; more ways to arrange→higher S
- State order at same T, p: gases $\gg$ liquids > solids; more particles and more volume usually → higher S
- **Qualitative ΔS predictions:**
- **Physical changes**: melting(+), vaporization(+), sublimation(+), freezing(−), condensation(−), deposition(−)
- **Particle number & gas moles**:
  - if $\Delta n(\text{gas})>0$ then ΔS usually >0
  - if $\Delta n(\text{gas})<0$ then ΔS usually <0
- **Mixing/dissolving**:
  - **mixing** different gases/solutions typically **increases** S;
  - **dissolving** ionic solids in water **often increases** S overall (ions dispersed) but can decrease if strong ordering of solvent dominates (rare at this level)
- **Standard entropy** calculation:

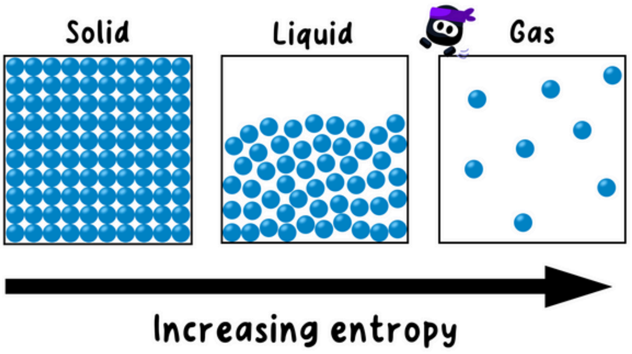
$$\Delta S^\ominus=\sum\left(S^\ominus \text{ products }\right)-\sum\left(S^\ominus \text{ reactants }\right)$$

Gibbs energy $\Delta G^\ominus=\Delta H^\ominus-T\Delta S^\ominus$ (HL)

- **Gibbs energy ΔG** measures maximum non-expansion work obtainable at constant T, p; combines **enthalpy** and **entropy effects**
- Formula use:
  - $\Delta G^\ominus(\text{kJ mol}^{-1}) = \Delta H^\ominus(\text{kJ mol}^{-1})-T(\text{K})\cdot\Delta S^\ominus(\text{kJ K}^{-1}\text{ mol}^{-1})$
  - convert $\Delta S^\ominus$ from J → kJ by ÷1000
- **Interpreting terms**:
  - $\Delta H^\ominus<0$ favors spontaneity
  - $\Delta S^\ominus>0$ favors spontaneity
  - T scales the entropy term

Common Mistake

Don't use °C instead of K in T and don't mix J and kJ, which leads to ×1000 errors.



ΔG away from equilibrium; Q and K relations (HL)

- **Nonstandard conditions**: $\Delta G=\Delta G^\ominus+RT\ln Q$ where Q is the reaction quotient (activities/approximate with partial pressures or concentrations for ideal cases)
- **At equilibrium**: $\Delta G=0$ and $Q=K$ → $\Delta G^\ominus=-RT\ln K$;
  - **large K** → $\Delta G^\ominus \ll 0$
  - **small K** → $\Delta G^\ominus \gg 0$
- **Constants & units**
  - $R=8.314\text{ J K}^{-1}\text{ mol}^{-1}$
  - T in K
  - ΔG in J mol^{-1} from RT ln terms (convert to kJ if needed)

Note!

Equilibrium constant K and reaction quotient Q are officially covered in **R2.3 How far? The extent of chemical change**.

Common Question

Which change has the greatest decrease in entropy?

A. $CO_2\,(s) \rightarrow CO_2\,(g)$

B. $CO_2\,(g) \rightarrow CO_2\,(l)$

C. $CO_2\,(g) \rightarrow CO_2\,(s)$

D. $CO_2\,(l) \rightarrow CO_2\,(s)$

Spontaneity, temperature thresholds, and ΔG sign (HL)

- **Spontaneity** rule: at constant p,T, the change is:
  - **spontaneous** if $\Delta G<0$
  - **non-spontaneous** if $\Delta G>0$
  - at **boundary** if $\Delta G=0$
- **Temperature of spontaneity:** Set $\Delta G^\ominus=0$→ $T=\Delta H^\ominus/\Delta S^\ominus$ (ensure consistent units; T in K)
- **System vs surroundings**
  - ΔG packs both **system** entropy change ($\Delta S(\text{sys})$) and **surroundings** entropy change via heat flow ($\Delta S(\text{sur})=-\Delta H/T$);
  - That's why $\Delta G=\Delta H-T\Delta S$ predicts **spontaneity**

ΔH	ΔS	ΔG	Spontaneity
< 0	> 0	Always negative	Always spontanous
> 0	< 0	Always positive	Always non-spontaneous
< 0	< 0	Negative at low T, positive at high T	Spontaneous at low T
> 0	> 0	Negative at high T, positive at low T	Spontaneous at high T

Common Mistake

Don't use K with nonstandard concentrations! Remember: $\Delta G^\ominus$ relates to **K at standard state**; otherwise use Q.

Access full IB Chemistry notes & practice questions on **revisiondojo.com**

IB CHEMISTRY R2.1 HOW MUCH? THE AMOUNT OF CHEMICAL CHANGE

Writing balanced equations with states

- Chemical equations encode **mole ratios** of reactants → products; **coefficients** are **simplest whole numbers**
- Include **state symbols**: (s), (l), (g), (aq)
- **Balance atoms first**, then charge (if ionic in aqueous), leave O_2/H_2O last when helpful

Example:

- Combustion of ethanol:
  $C_2H_5OH(l)+3O_2(g) \rightarrow 2CO_2(g)+3H_2O(l)$
- Ionic precipitation:
  $AgNO_3(aq)+NaCl(aq) \rightarrow AgCl(s)+NaNO_3(aq)$

Using mole ratios to get masses, volumes, concentrations

- **Core links**
  - $n = m/M$
  - $n = C \cdot V(dm^3)$;
  - for gases at same T, p: $V \propto n$
  - at STP $Vm \approx 22.7\ dm^3\ mol^{-1}$
- **Procedure**
  - Balance equation
  - Identify known → moles
  - Use mole ratio to find unknown moles
  - Convert to requested unit

Example:

What **mass** of $CaCO_3$ decomposes to make 4.40 g CO_2?

$$CaCO_3(s) \rightarrow CaO(s)+CO_2(g)$$
$$n(CO_2) = 4.40/44.01 = 0.100 \text{ mol}$$
$$n(CO_2):n(CaCO_3) = 1:1 \rightarrow n(CaCO_3) = 0.100 \text{ mol}$$
$$\rightarrow m = 0.100 \times 100.09 = 10.01 \text{ g}$$

Limiting reactant: theoretical yield vs experimental yield

- **Limiting reactant (LR)**: consumed first; fixes maximum amount of product
- **Theoretical yield**: amount of product predicted from LR via stoichiometry
- **Experimental/actual yield**: amount actually obtained in lab

Exam Tip!

To **identify LR**, convert each reactant amount to moles of the same product: the smaller result indicates the LR.

Percentage yield

$$\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

- **Reasons for <100%**:
  - side reactions
  - incomplete reaction
  - loss on transfer/purification
  - impure reagents
- **Reasons for >100%**:
  - wet/impure product
  - solvent trapped
  - calibration errors

Example:

- Assume you are given theoretical NH_3 = 12.2 g
- If 10.0 g collected
  → **% yield** = (10.0/12.2)×100 = 82.0%

Atom economy

$$\text{Atom Economy} = \frac{\text{Molar Mass of Desired Product}}{\text{Sum of Molar Masses of All Products}} \times 100$$

- **Stoichiometric**, **theoretical metric:** does not include yield or process energy/solvent impacts
- **Addition/rearrangement reactions** often have high (≈100%), **substitutions/eliminations** - lower

Example:

Production of ethene by dehydration:

$$C_2H_5OH \rightarrow C_2H_4+H_2O \text{ (desired } C_2H_4)$$

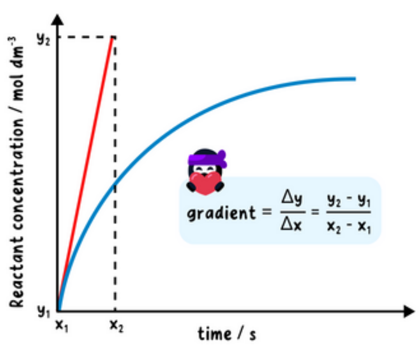
- M(desired) = 28.05;
- ΣM(products) = 28.05+18.02 = 46.07

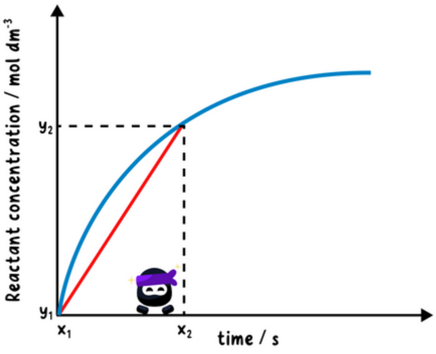
→ **Atom economy** = (28.05/46.07)×100 = 60.9%

IB CHEMISTRY R2.2 HOW FAST? THE RATE OF CHEMICAL CHANGE

Defining and determining reaction rate

- **Rate** = ±Δ[species]/Δt; use product rate positive, reactant rate positive by convention with a minus sign
- **Units**:
  - mol dm^{-3} s^{-1} (or M s^{-1})
  - gas volume data can be converted to moles via molar volume or PV=nRT
- **From graphs**:
  - **Instantaneous rate** = slope of tangent to concentration (or volume/mass) vs time curve at a chosen t
  - **Average rate** = secant line connecting two points on the curve





Exam Tip!

To **draw tangent lines accurately**, use a ruler and ensure the line touches the curve at only one point without crossing it.

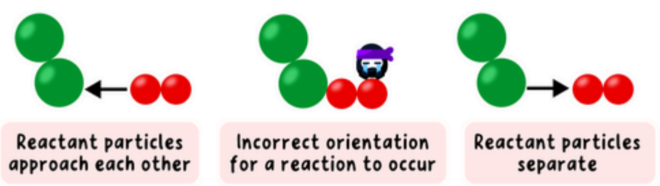
Collision theory

- **Kinetic energy vs temperature**: average Ek ∝ T(K)
- **Higher T** → particles **move faster on average** and collide more often
- **Sufficient energy**: only collisions with **E≥Ea** can react
- **Proper orientation**: reactants must collide with a geometry that allows bonds to break/form (steric factor)
- **Why T matters twice**: higher fraction of molecules above Ea and higher collision frequency

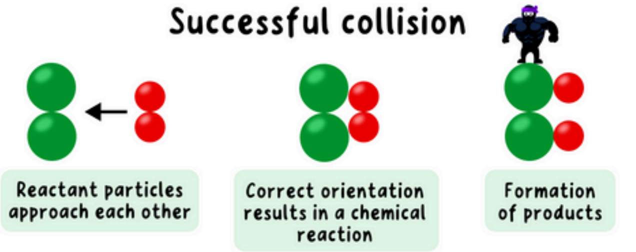
Exam Tip!

When you explain a **rate increase with T**, always mention "fraction of collisions with E≥Ea increases" and "more frequent collisions" (orientation unchanged)

Unsuccessful collision

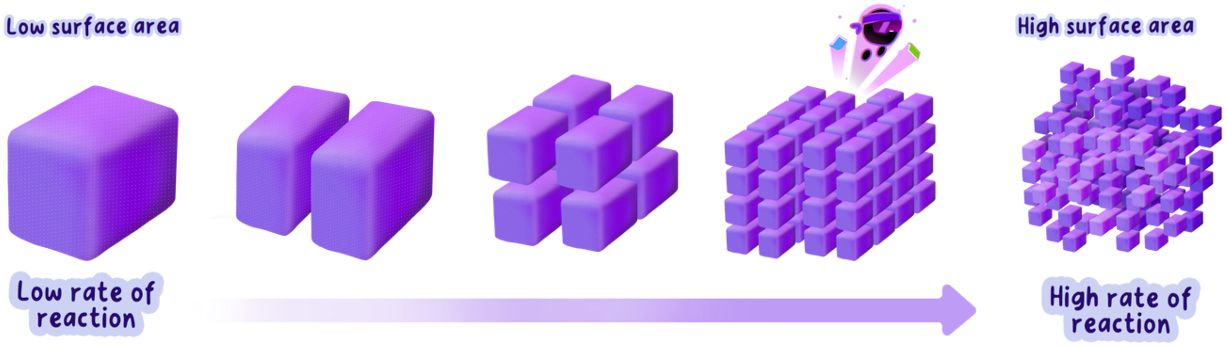


Successful collision



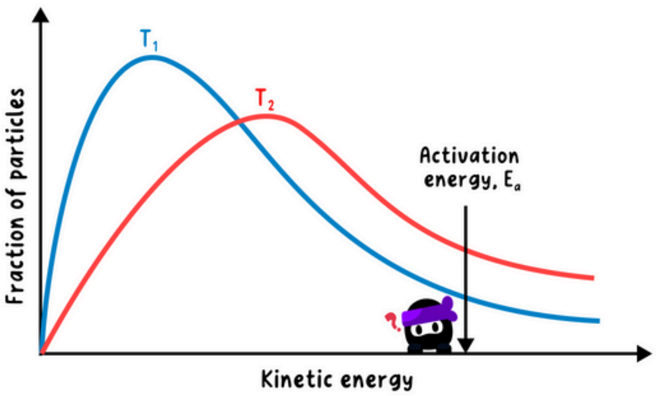
Factors affecting reaction rate

- **Concentration (liquids/solutions)** ↑ → more particles per volume → collision frequency ↑ → **rate** ↑
- **Pressure (gases)** ↑ at constant T → effective concentration ↑ → **rate** ↑
- **Surface area (solid reactant)** ↑ → more exposed sites → collision frequency at surface ↑ → **rate** ↑
- **Temperature** ↑ → fraction with E≥Ea ↑ and collisions ↑ → **rate** ↑
- **Catalyst** → alternative pathway with lower Ea → greater fraction of effective collisions at same T → **rate** ↑



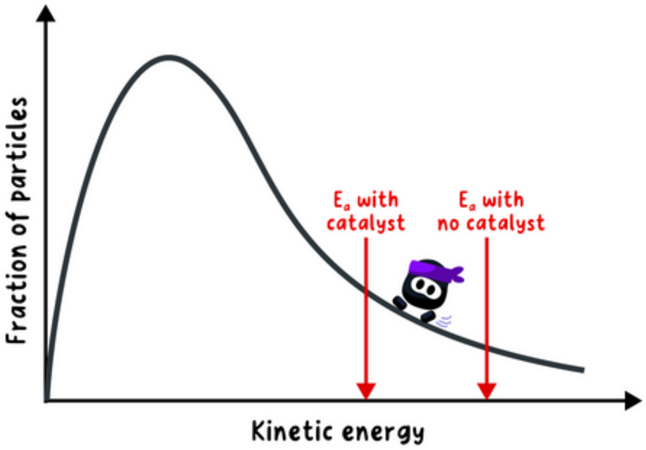
Activation energy & Maxwell-Boltzmann (M-B) distributions

- **M-B curve features**:
  - area under curve = number of particles(constant)
  - peak at most probable energy
  - tail extends to high E
- **Effect of increasing T**:
  - peak **lowers** and **shifts right**
  - **area to the right of Ea** (between the curve and horizontal axis) greatly **increases**



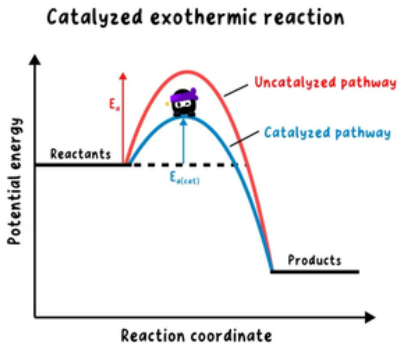
Catalysts: energy profiles and M-B distributions

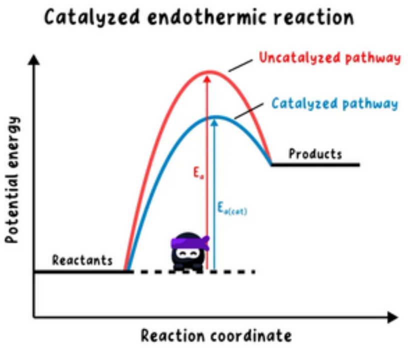
- **Energy profiles (exothermic)**:
  - y = potential energy
  - x = reaction coordinate
  - uncatalyzed has higher hump
  - catalyzed path has lower hump(s);
  - ΔH same for both
- **Energy profiles (endothermic)**:
  - same idea but products above reactants
  - catalyst reduces Ea (forward and reverse) equally
  - ΔH unchanged
- **M-B with catalyst**:
  - draw one temperature curve and two vertical lines Ea(uncat) and Ea(cat) (left-shifted)
  - shaded "above Ea" region is larger for catalyzed path

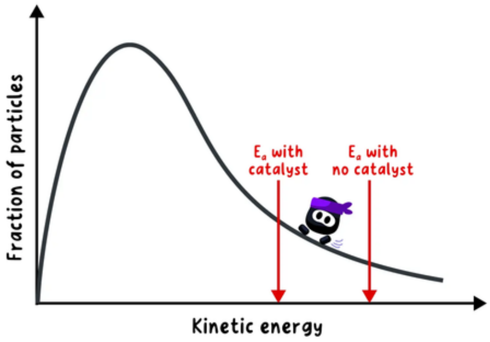


Common Mistake

Don't claim that catalysts "increase the number of collisions"! They primarily **increase the fraction of collisions** that are successful by lowering Ea (though heterogeneous catalysts can also increase effective collision frequency at active sites).

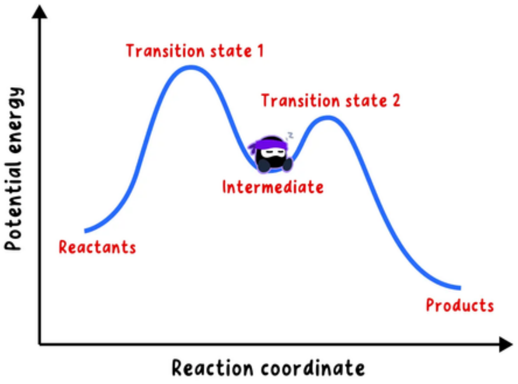






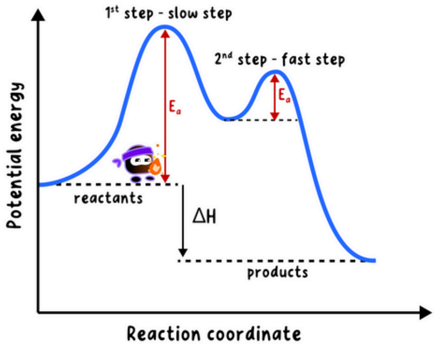
Mechanisms, intermediates, transition states

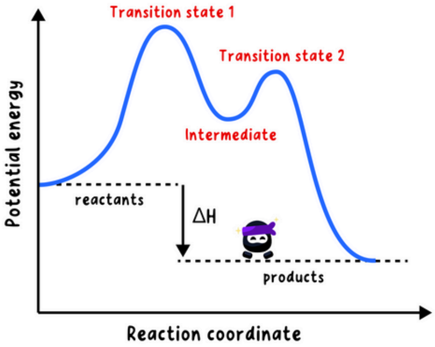
- **Mechanism** = sequence of elementary steps that sum to overall reaction; must match overall stoichiometry and kinetics
- **Rate-determining step(RDS)** = slowest elementary step; controls the rate law form
- **Intermediate** = species formed then consumed; appears in mechanism steps, not in overall equation; can sometimes be detected/trapped
- **Transition state** = highest-energy configuration along a step; not isolable; shown as peaks on energy profile



Multistep energy profiles & identifying the RDS (HL)

- **Energy profile axes**:
  - x = reaction coordinate
  - y = potential energy
- **Multistep** = multiple peaks(transition states);
- **Valleys** = intermediates
- **RDS** = step with tallest peak(highest Eaa); its transition state controls the rate
- **How to construct from data**: more steps → more distinct kinetic features (e.g., induction period suggests intermediate buildup); catalyst lowers one or more peaks but leaves ΔH unchanged





Common Mistake

Don't call the highest intermediate "transition state": **peaks** = transition state , **valleys** = intermediates.

Molecularity of elementary steps (HL)

- **Unimolecular**:
  - one reacting particle in step(A→products)
  - typical of isomerizations/decompositions
- **Bimolecular**:
  - two particles collide(A+B→… or A+A→…)
  - most common
- **Termolecular**:
  - three particles in one step(A+B+C→…)
  - rare(improbable simultaneous collision)

Rate equations (HL)

- **General form**: rate=k[A]^m[B]^n…; m,n are orders found experimentally
- **Method (slope or initial rates)**: hold all but one concentration constant; vary one and measure initial rate
- **Quick inference rules**:
  - if doubling [A] doubles rate → first order in A
  - if rate unchanged → zero order
  - if rate quadruples → second order

Molecularity of elementary steps (HL)

- **Order wrt a reactant** = exponent in rate law;
- **Overall order** = sum of exponents
- **Connection to mechanism**: for RDS that is elementary, orders match molecularity of that step

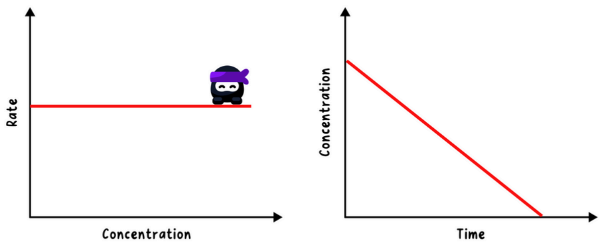
Exam Tip!

State axes explicitly; annotate slope/intercept relation to k on the chosen linearized plot

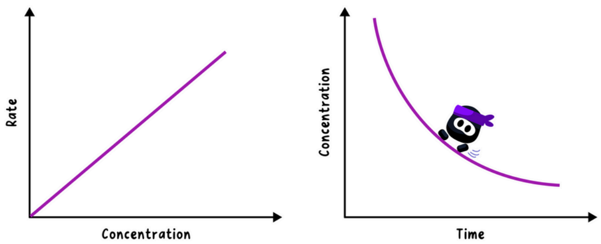
Characteristic concentration-time([A] vs t):
- Zero-order: linear decay (slope=−k)
- First-order: exponential decay; straight line in ln[A] vs t with slope = −k
- Second-order (1:1): curved; straight line in 1/[A] vs t with slope = +k

Characteristic rate-concentration(rate vs [A]):
- Zero: horizontal line (rate=k)
- First: straight line through origin (slope=k)
- Second: parabola (rate=k[A]^2)

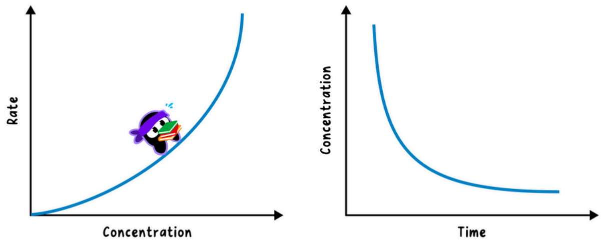
Zero-order reaction



First-order reaction

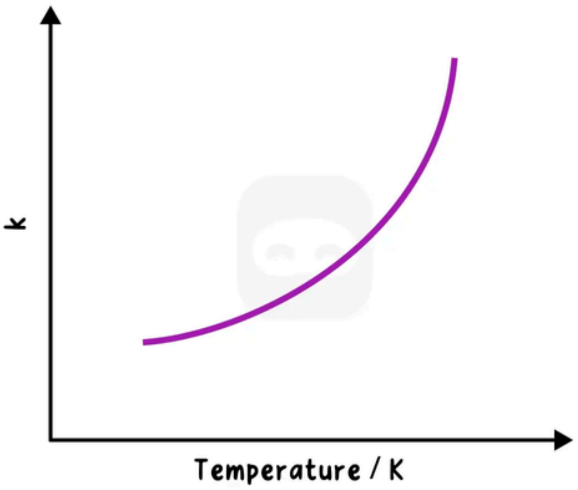


Second-order reaction



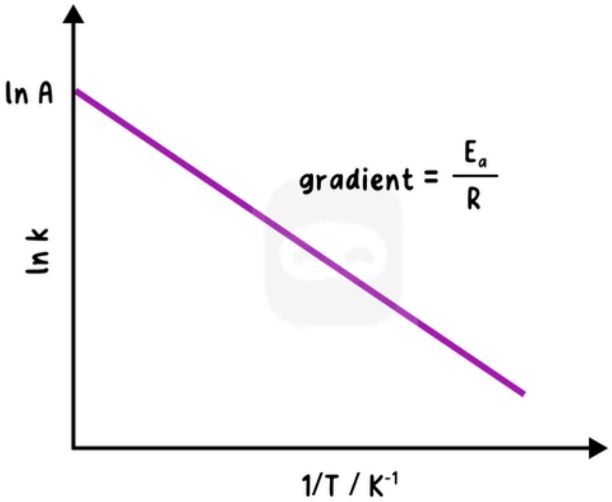
Rate constant (HL)

- **Units** depend on **overall order** so that rate (mol dm^{-3} s^{-1}) = k·(concentration)^order
- **Common unit patterns**:
    - Overall **zero** → k: mol dm^{-3} s^{-1}
    - Overall **first** → k: s^{-1}
    - Overall **second** (1:1) → k: dm^3 mol^{-1} s^{-1}
    - Overall **third** → k: dm^6 mol^{-2} s^{-1}
- Procedure: solve k = rate/([A]^m[B]^n…) with SI-consistent mol dm^{-3} and s



Arrhenius equation (HL)

- **Qualitative link**: as T↑, k↑ because a larger fraction of molecules have E≥Ea and A can also be T-dependent slightly
- **Arrhenius forms**:
    - $k = A \cdot e^{(-Ea/RT)}$
    - $\ln k = \ln A - (Ea/R)(1/T)$ → straight line: $y = \ln k$, $x = 1/T$, slope = $-Ea/R$, intercept = $\ln A$
- Using two temperatures: $\ln(k_2/k_1) = (Ea/R)(1/T_1 - 1/T_2)$
- Graph reading:plot ln k vs 1/T; find slope m; Ea = −mR; A = e^(intercept)



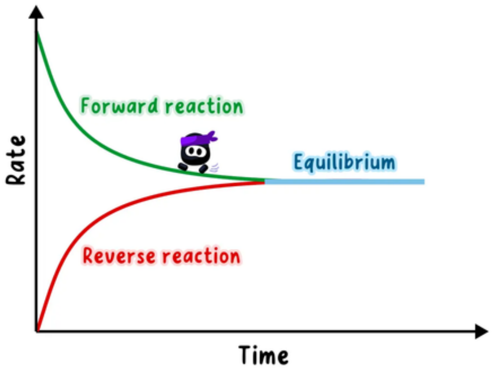
Arrhenius factor A (HL)

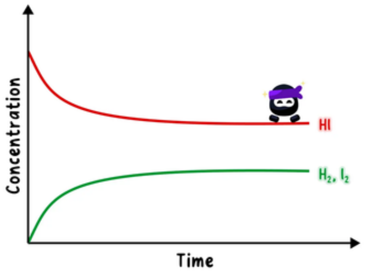
- **Physical meaning of A**:
    - frequency of properly oriented, collision-ready encounters(“pre-exponential”)
    - includes steric factor and collision frequency

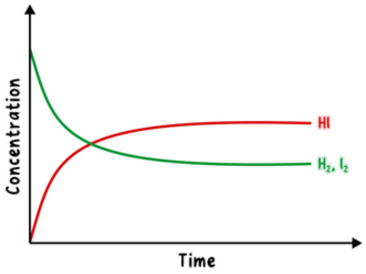
IB CHEMISTRY R2.3 HOW FAR? THE EXTENT OF CHEMICAL CHANGE

Dynamic equilibrium in closed systems

- **Definition**:
  - in a closed system, equilibrium is reached when rate(forward) = rate(backward);
  - macroscopic properties (concentration, pressure, color) are constant in time.
- **Characteristics**:
  - Both reactions continue (dynamic, not static).
  - Concentrations are constant but not necessarily equal.
  - Equilibrium composition depends on T (and total pressure for gaseous systems).
  - Reached from either side (same final composition at given T, total n, V).
- **Physical equilibria**: e.g., $H_2O(l) \rightleftharpoons H_2O(g)$ in a sealed flask (evaporation = condensation).
- **Chemical equilibria**: e.g., $N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$ (colorless ⇌ brown).







The equilibrium law & constant

- For a homogeneous reaction: $aA+bB \rightleftharpoons cC+dD$

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

- Use activities ≈ concentrations (mol dm^{-3}) for solutions; for gases, a Kp may be used but is not required by IB syllabus.
- **Omit** pure solids and pure liquids (activity ≈ 1).
- Always include **stoichiometric powers**.

Le Châtelier's principle

Predict **qualitative shifts** when a stress is applied to a system at equilibrium:

- **Concentration**:
  - add a reactant → shifts right
  - remove it → left. K unchanged.
- **Pressure/volume (gases)**:
  - decrease V/increase P → shifts to fewer moles of gas
  - increase V → more moles of gas. K unchanged.
- **Temperature**: treat heat as a reagent.
  - **Exothermic** ($\Delta H<0$): ↑T shifts left and decreases K.
  - **Endothermic** ($\Delta H>0$): ↑T shifts right and increases K.

Common Mistake

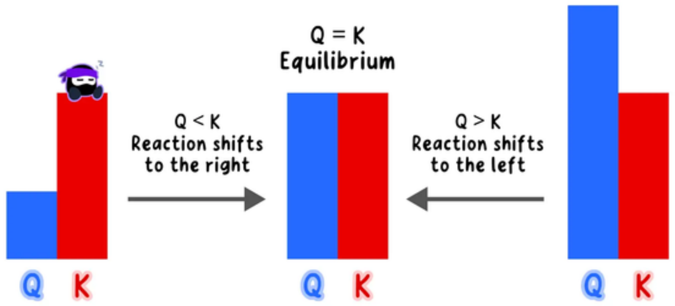
Don't claim that **catalysts** change K. They don't; they change rate to reach equilibrium, not the position.

Magnitude of K & temperature dependence

- **Extent at fixed T**:
  - $K \ll 1$: mostly reactants.
  - $K<1$: reactants favored.
  - $K=1$: appreciable amounts of both.
  - $K>1$: products favored.
  - $K \gg 1$: mostly products.
- **Relationships at the same T**:
  - **Reverse** reaction: $K_{rev}=1/K$.
  - **Multiplying** coefficients by n: $K'=K^n$.
  - **Adding** reactions: $K_{sum}=K_1 \times K_2$ (multiply constants).

Reaction quotient (HL)

- Use the same expression as K but with instantaneous concentrations/pressures:
  - **If Q<K**: system has too many reactants → shifts right to make products.
  - **If Q>K**: system has too many products → shifts left to make reactants.
  - **If Q=K**: at equilibrium.



Quantifying equilibrium compositions (HL)

Workflow (homogeneous only):

- Write K expression.
- Set up **I-C-E (Initial-Change-Equilibrium) table** in molarity.
- Express equilibrium concentrations in terms of x; substitute into K.
- If K is **very small** or **large**, use the small-x approximation (e.g., $a-x \approx a$) and check after.
- **Approximation cue**: if the computed $x/a < 5\%$, your approximation is fine.

Gibbs free energy and equilibrium (HL)

Relationship (data booklet):

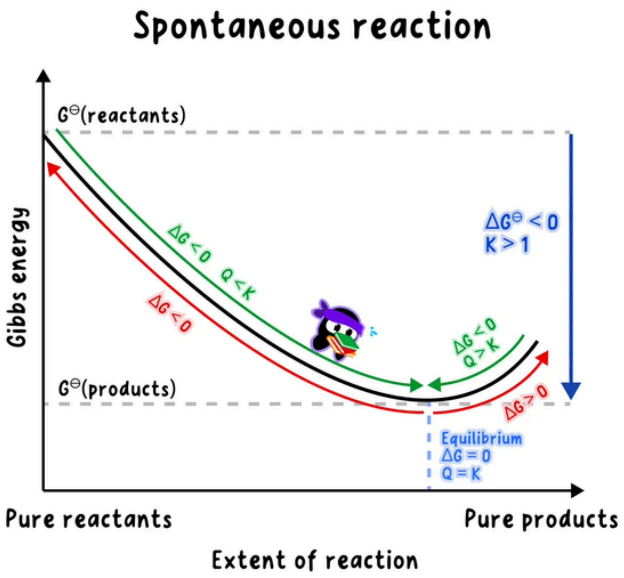
$$\Delta G^\circ = -RT \ln K$$

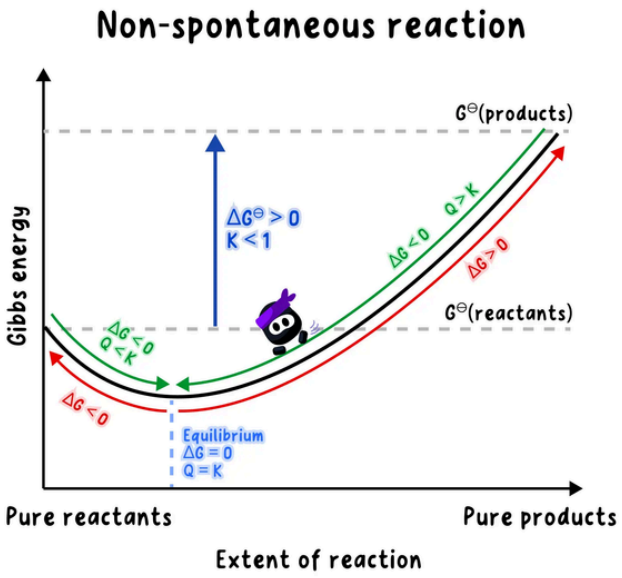
At **fixed T**:

- $K>1 \Rightarrow \Delta G^{\ominus}<0$ (products favored).
- $K=1 \Rightarrow \Delta G^{\ominus}=0$
- $K<1 \Rightarrow \Delta G^{\ominus}>0$ (reactants favored).

Common Mistake

Don't mix J and kJ! Keep track of that when using **R**.





IB CHEMISTRY R3.1 PROTON TRANSFER REACTIONS

Brønsted-Lowry acids and bases

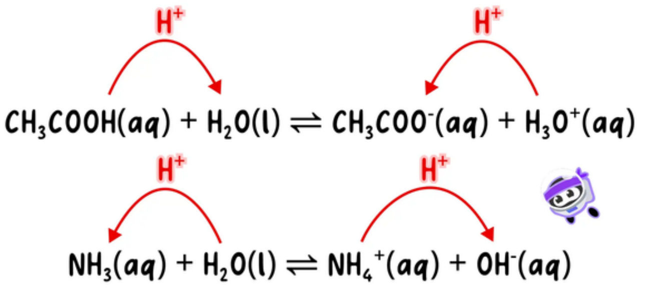
- **Acid** = proton (H^+) donor
- **Base** = proton acceptor.
- In water, write the proton as $H^+(aq)$ or $H_3O^+(aq)$.
- **Base vs alkali:** a base accepts H^+; an alkali is a soluble base that produces OH^- in water (e.g., NaOH, KOH).
- **How to identify in an equation**: Look for the species losing H^+ (acid) and the one gaining H^+ (base) in the same step.

Example:

- $HCl + H_2O \rightarrow H_3O^+ + Cl^-$
  - → HCl acid (donor), H_2O base (acceptor).
- -$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
  - → NH_3 base, H_2O acid.
- -$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$
  - → ethanoic acid donates to water.

Amphiprotic species

- Identify species that can **both donate** and **accept H^+**: H_2O, HCO_3^-, HSO_4^-, HS^-, $H_2PO_4^-$.
- Write **both reactions**:
  - $HCO_3^- + H_3O^+ \rightleftharpoons H_2CO_3 + H_2O$ (H_2O acts as base)
  - $HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+$ (H_2O acts as acid)



Ion product of water Kw

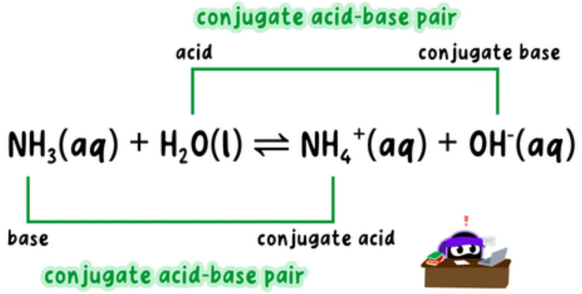
- At 298 K, Kw = 1.0×10^{-14}.
- Neutral water: $[H^+] = [OH^-] = 1.0\times10^{-7}$ M → pH = 7.00.
- If $[H^+] > [OH^-]$ → **acidic**
- If $[H^+] < [OH^-]$ → **basic**
- Use $[OH^-] = K_w/[H^+]$ and $pOH = -\log[OH^-]$; pH + pOH = 14.00 (298 K).

Common Mistake

Don't forget about the effect of **temperature**: pH 7 is neutral only at 298 K.

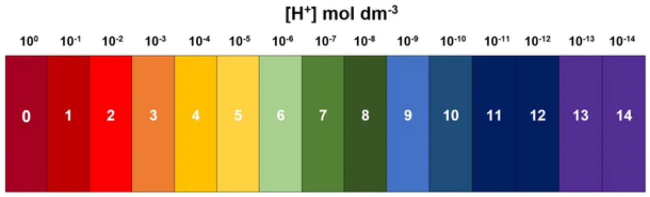
Conjugate acid-base pair

- **Rule**:
  - conjugate base = acid minus H^+;
  - conjugate acid = base plus H^+.
  - **Pairs differ by one proton**.
- **Quick conversions**
  - Base → conj. acid: add H^+ and +1 charge.
  - Acid → conj. base: remove H^+ and −1 charge.



pH and $[H^+]$

- **Formulas (data booklet)**:
  - $pH = -\log_{10}[H^+]$
  - $[H^+] = 10^{(-pH)}$ (units: mol dm^{-3})
- **Estimating/measuring pH**: universal indicator (rough), pH meter/probe (precise).



Strong vs weak acids and bases

- **Strong acids** (≈complete ionization in water): HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4 (first H^+).
- **Strong bases**: Group 1 hydroxides (NaOH, KOH, etc.), $Ba(OH)_2$ (highly soluble).
- **Weak acids/bases**: partial ionization (e.g., CH_3COOH, H_2CO_3, NH_3, amines).
- **Equilibria favor** the **weaker conjugate pair** (e.g., strong acid has a very weak conjugate base).

Note!

- **Strength** = degree of ionization at a given concentration.
- **Concentration** = amount dissolved (mol dm^{-3}). A strong acid can be dilute; a weak acid can be concentrated.

Neutralization

General reactions (don't forget writing states):

- Acid + metal oxide → salt + water
- Acid + metal hydroxide → salt + water
- Acid + carbonate → salt + CO_2 + H_2O
- Acid + hydrogencarbonate → salt + CO_2 + H_2O
- Acid + ammonia/amine → ammonium/alkylammonium salt (no H_2O unless extra base present)

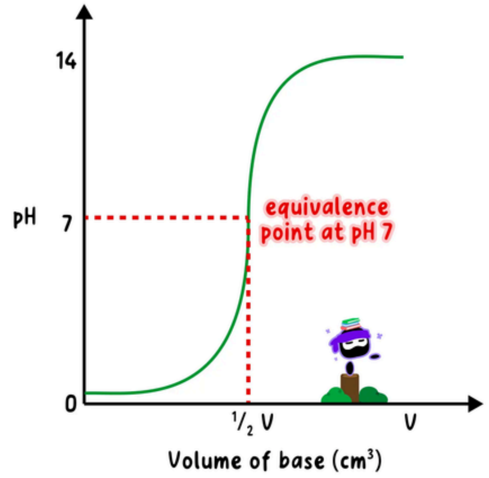
Exam Tip!

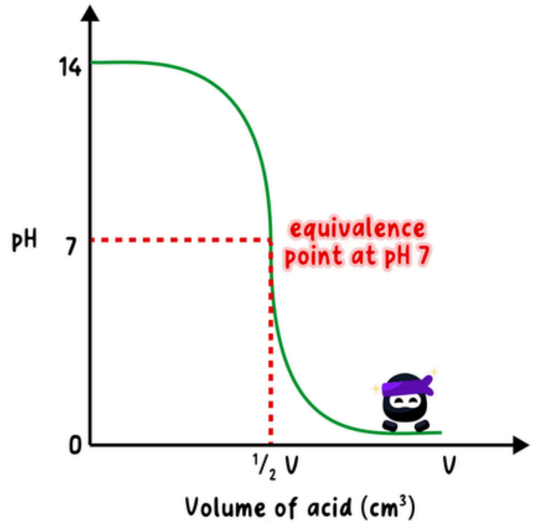
For **carbonates**, always show $CO_2(g)$ + $H_2O(l)$.

pH curves for neutralization

General features (titrant added on x-axis; pH on y-axis):

- **Strong acid + strong base**:
  - initial pH low
  - sharp vertical rise near equivalence
  - equivalence at pH ≈ 7
- **Strong base + strong acid**:
  - initial pH high
  - sharp vertical fall near equivalence
  - equivalence at pH ≈ 7
- **Intercepts & equivalence**
  - Initial pH from analyte strength
  - Equivalence point where mol H^+ = mol OH^- (steepest slope).
- **Choice of indicator**: pick one whose range overlaps the steep vertical region.

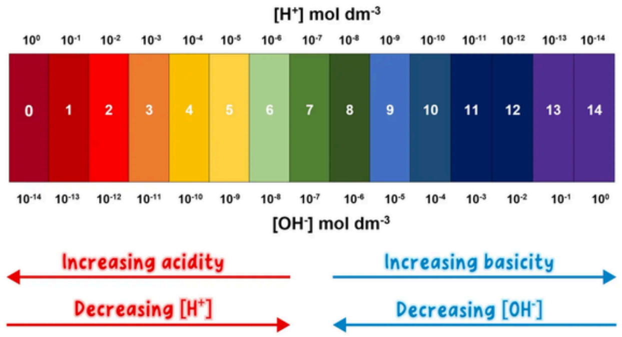




Quantifying equilibrium compositions (HL)

Definitions (298 K unless stated):

- **pOH** = $-\log_{10}[OH^-]$
- **$[OH^-]$** = 10^(−pOH)
- **pH + pOH = 14.00** (because $K_v = [H^+][OH^-] = 1.0\times10^{-14}$ at 298 K).
- **Interconvert quickly**: find one of {pH, pOH, $[H^+]$, $[OH^-]$} → get the other three via the two equations above.



Strength of weak acids and bases (HL)

$$K_a = \frac{[H^+][A^-]}{[HA]} \qquad K_b = \frac{[BH^+][OH^-]}{[B]}$$

- **Trend sense**:
  - larger Ka / smaller pKa → stronger acid
  - larger Kb / smaller pKb → stronger base
- **Comparisons**:
  - If $pKa(HA_1) < pKa(HA_2)$ → HA_1 is stronger
  - If $pKb(B_1) < pKb(B_2)$ → B_1 is stronger

Conjugate acid-base pairs (HL)

- For $HA \rightleftharpoons H^+ + A^-$ and $A^- + H_2O \rightleftharpoons HA + OH^-$, multiplying Ka and Kb cancels intermediates → Ka·Kb = Kw (at same T).
- **Uses**:
  - Given Ka, find Kb of its conjugate base: Kb = Kw/Ka (and vice versa).
  - Convert pKa ↔ pKb: pKa + pKb = pKw (= 14.00 at 298 K).

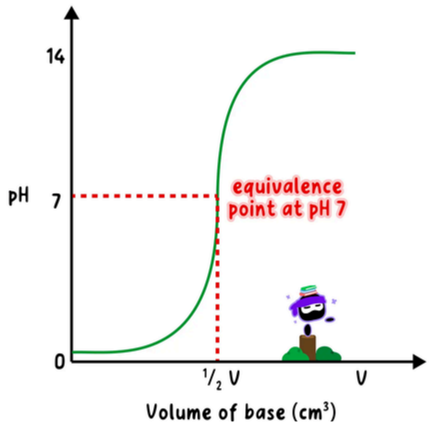
The pH of a salt solution (HL)

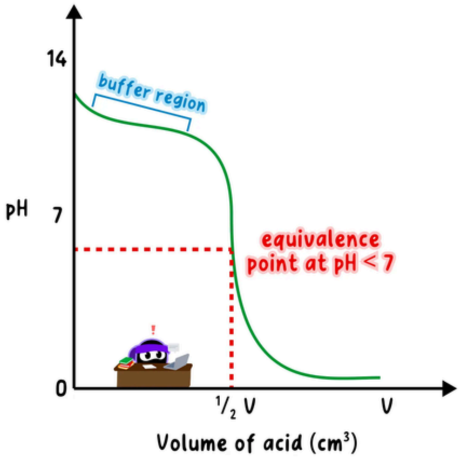
- Decide if **acidic/basic/neutral salt** from parent acid/base strengths:
  - Strong acid + strong base → neutral salt (e.g., NaCl).
  - Strong acid + weak base → acidic salt (e.g., NH_4Cl).
  - Weak acid + strong base → basic salt (e.g., CH_3COONa).
  - Weak acid + weak base → pH depends on relative Ka vs Kb.
- **Hydrolysis equations** (write the reacting ion with water):
  - $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ (acidic)
  - $RCOO^- + H_2O \rightleftharpoons RCOOH + OH^-$ (basic)
  - $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$ (basic)
  - HCO_3^- amphiprotic:
    - as base: $HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$
    - as acid: $HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+$

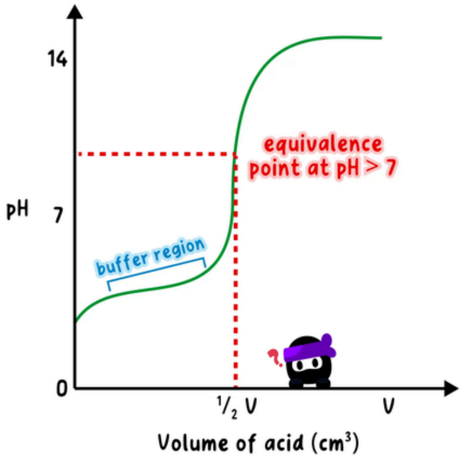
pH curves of acid-base combinations (HL)

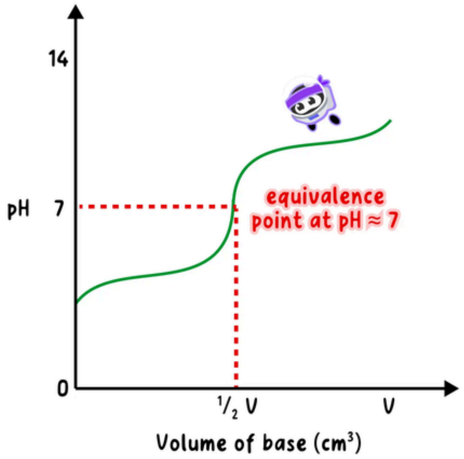
General shapes & landmarks (titrant on x-axis):

- **Strong acid + strong base**: start low pH; equivalence ≈ 7; very steep near eq.
- **Weak acid + strong base**: higher initial pH; buffer region; ½-equivalence: pH = pKa; equivalence > 7.
- **Strong acid + weak base**: lower initial pH than pure base; buffer region; ½-equivalence: pOH = pKb; equivalence < 7.
- **Weak acid + weak base**: shorter vertical; equivalence pH depends on pKa vs pKb; indicator choice is tricky.









Acid-base indicators and pH changes (HL)

- **Model**: HInd ⇌ H^+ + Ind^-, where HInd and Ind^- are different colors.
- If **pH ≪ pK(Ind)** → HInd color dominates.
- If **pH ≫ pK(Ind)** → Ind^- color dominates.
- **End point** ≈ pK(Ind) (mid-transition).
- **Universal indicator** = mixture → wide, multicolor range.

Common Mistake

Don't equate end point with equivalence point, because they should coincide only when the indicator is well chosen.

Choosing an indicator (HL)

- **Rule**: pick an indicator whose transition range overlaps the steep pH jump at equivalence.
- **Equivalence point** = stoichiometric H^+=OH^-
- **End point** = observed color change (≈ indicator pKa).

Common Question

Which conditions best favour the oxidation of secondary alcohols to ketones?

A. Excess acidified potassium dichromate (VI) and reflux

B. Excess acidified potassium dichromate (VI) and distillation

C. Few drops of acidified potassium dichromate (VI) and reflux

D. Few drops of acidified potassium dichromate (VI) and distillation

Buffer solutions (HL)

- **Acidic buffer**: weak acid HA + its salt A^- (e.g., CH_3COOH/CH_3COONa).
- **Basic buffer**: weak base B + its salt BH^+ (e.g., NH_3/NH_4Cl).
- **How buffers resist pH change**:
  - Add acid (H^+): A^- consumes H^+ → HA forms.
  - Add base (OH^-): HA consumes OH^- → A^- + H_2O forms.
- **Buffer capacity** increases with total concentration and with [A⁻]≈[HA] (best near pH≈pKa).

pH of buffer solutions (HL)

- **Acidic buffer**:

$$K_a = \frac{[H^+][A^-]}{[HA]} \Rightarrow pH = pK_a + \log\frac{[A^-]}{[HA]}$$

- **Basic buffer**:

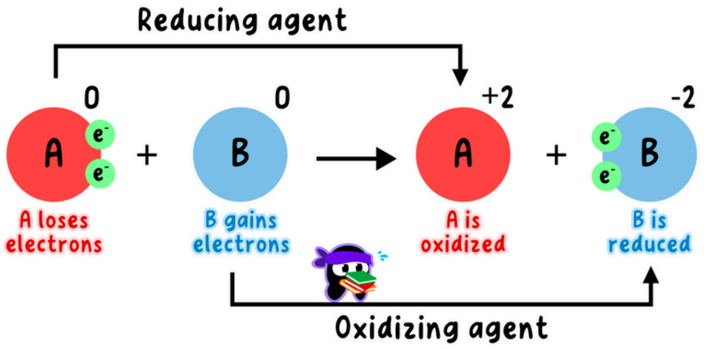
$$pOH = pK_b + \log\frac{[BH^+]}{[B]}, \quad pH = 14.00 - pOH(298\ \mathrm{K})$$

- **Effect of dilution**: if both components diluted by the same factor, the ratio [A−]/[HA] stays ≈ constant → pH ~ unchanged, but buffer capacity decreases (less moles to neutralize added acid/base).

IB CHEMISTRY R3.2 ELECTRON TRANSFER REACTIONS

Neutralization

- **Redox lenses**:
    - **Electron**: oxidation = e^- loss; reduction = e^- gain.
    - **O state**: oxidation = increase in oxidation state; reduction = decrease.
    - **O/H**: oxidation = O gain / H loss; reduction = O loss / H gain.
- **Assigning oxidation states** (quick rules): free element = 0; monoatomic ion = charge; Group 1 = +1, Group 2 = +2; F = −1; O = −2 (peroxide −1); H = +1 (metal hydrides −1); sums → 0 (molecule) or ion charge.
- **Identify roles in an equation**: the species oxidized (OS ↑) is the reducing agent; the species reduced (OS ↓) is the oxidizing agent.
- **Variable OS examples**: Fe^{2+}/Fe^{3+}; Mn from +2 to +7 (e.g., MnO_4^-); S (−2, 0, +4, +6); N (−3 to +5). Names use Roman numerals, e.g., iron(III) oxide.



Neutralization

- Write **separate oxidation** and **reduction half-equations**; balance atoms, then charge with e^-; add to cancel e^-.
- **Acidic solution method**: balance O with H_2O, H with H^+, then charge with e^-.
- **Neutral (or basic*) shortcut for IB SL**: use the acidic method then (if basic) add same number of OH^- both sides to neutralize H^+ to water.

Exam Tip!

Ensure **electron count** cancels exactly, and then remove any spectator species.

Reactivity of elements based on periodic trends

- **Metals**:
    - ***across a period*** (to the right) metals become less easily oxidized
    - ***down Group 1/2***, larger atoms with lower IE are more easily oxidized.
    - A more reactive metal displaces a less reactive metal from its ion solution.
- **Halogens**: oxidizing power decreases down Group 17; $F_2 > Cl_2 > Br_2 > I_2$ (so Cl_2 oxidizes Br^-/I^-; Br_2 oxidizes I^-).
- **Using supplied data**: if $M(s) + M'^{n+}(aq) \rightarrow$ reaction occurs, M is the stronger reducing agent (easier to oxidize).

Acid-metal reactions

- **General**: metal + dilute $HCl/H_2SO_4 \rightarrow$ salt + H_2 (metals above H in the activity series).

Example:

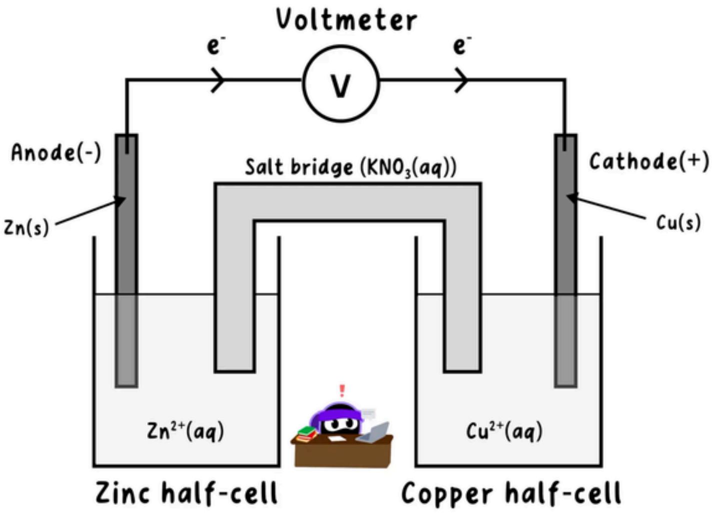
- $Mg + 2HCl \rightarrow MgCl_2 + H_2\uparrow$
- $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$
- $Fe + 2HCl \rightarrow FeCl_2 + H_2\uparrow$

Primary (voltaic) cell

- **Build**: two half-cells (e.g., $Zn|Zn^{2+}$ and $Cu|Cu^{2+}$), metal electrodes, connecting wire + voltmeter, salt bridge ($KNO_3(aq)$) to maintain charge balance.
- **Flow**: electrons from anode (oxidation) to cathode (reduction).
- In **salt bridge**, anions drift toward anode compartment; cations toward cathode.

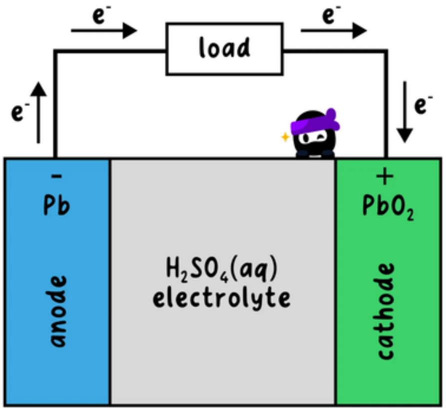
Electrochemical cells: oxidation and reduction

- **Oxidation** at anode, **reduction** at cathode **(always!)**.
- **Voltaic** (galvanic, spontaneous): anode is −, cathode is +; e^- flow anode → cathode through wire.
- **Electrolytic** (non-spontaneous, power supply): anode is +, cathode is − (polarity set by source).



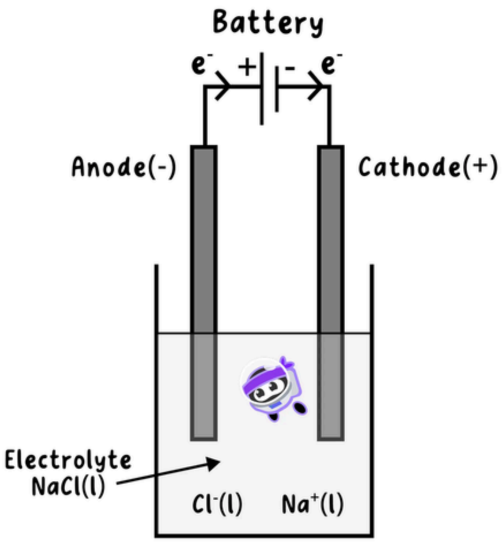
Secondary (rechargeable) cells

- **Secondary cells**: discharge reactions are reversed on charging by applying external voltage > E(cell).
- Given **discharge half-reactions**, reverse them (swap anode/cathode) for charge.
- **Pros/cons (snapshot)**:
    - **Primary**: convenient, high energy density, single-use waste.
    - **Secondary**: rechargeable, lower waste, finite cycle life, charging losses.
    - **Fuel cells**: continuous fuel → electricity, high efficiency/low local emissions; need fuel supply & infrastructure.



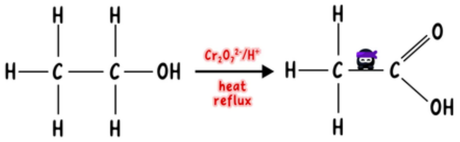
Electrolytic cells

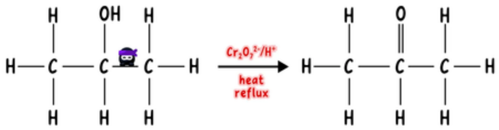
- **Conduction**:
    - **electrons** move in the external circuit
    - **ions** carry charge in the molten/aqueous electrolyte
    - **power source** forces non-spontaneous redox.
- **Products (molten salts, no water present)**:
    - **cations** → cathode (reduced to metal)
    - **anions** → anode (oxidized to non-metal).

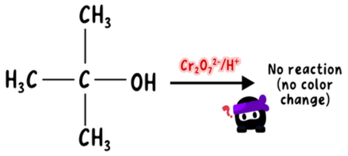


Oxidation of organic compounds

- **Primary alcohols** (1°): two-step oxidation
    - 1° alcohol → aldehyde (partial oxidation) → carboxylic acid (further oxidation); $R\text{-}CH_2\text{-}OH \rightarrow R\text{-}CHO \rightarrow R\text{-}COOH$
- **Secondary alcohols** (2°): $R_1\text{-}CH(OH)\text{-}R_2 \rightarrow$ ketone $R_1\text{-}C(=O)\text{-}R_2$
- **Tertiary alcohols** (3°): no reaction under these conditions (no H on the -C-OH carbon).
- **Setups**:
    - **Distillation** to isolate aldehyde: heat gently with oxidant; aldehyde distilled off as it forms.
    - **Reflux** to drive to acid: heat under condenser so vapors return; maximizes contact with oxidant.





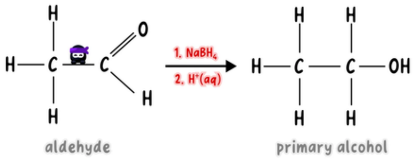


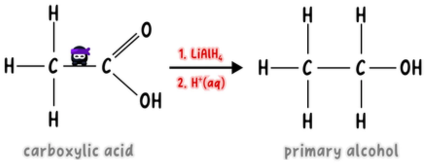
Reduction of functional groups

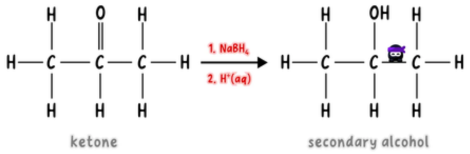
- **Carboxylic acid** → aldehyde → 1° alcohol (overall 4 H added: two to C=O, two to O-H stage).
- **Ketone** → 2° alcohol.
- **Role of hydride (H^-)**: delivers electrons + H to the electrophilic carbonyl carbon (you may write "+2[H]" per reduction step if reagents are not named).

Common Mistake

Don't reduce a **3° alcohol** to a ketone: it is impossible without oxidation; reductions go toward higher H / lower O at carbon.

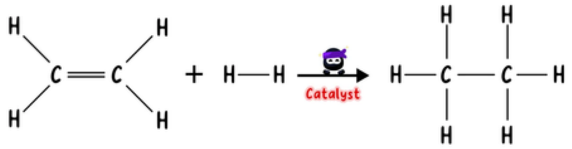






Hydrogenation of unsaturated compounds

- **Alkenes**: R-CH=CH-R′ + H_2 → R-CH_2-CH_2-R′ (alkane).
- **Alkynes**: R-C≡C-R′ + 2 H_2 → R-CH_2-CH_2-R′ (full hydrogenation).
- **"Reduction" view**: C gains H (decrease in oxidation state).



Standard cell potential & spontaneity (HL)

- Compute E°cell = E°(cathode, red) – E°(anode, red).
- If **E°cell > 0**, **forward** reaction is **spontaneous** (under standard conditions)
- If **E°cell < 0**, **reverse** is **spontaneous**.

Electrolysis of aqueous solutions (HL)

General rule of thumb (use E° and context):

- **Cathode (reduction)**:
  - If aqueous NaCl: H_2O/H_2 is typically reduced ($2\ H_2O + 2\ e^- \rightarrow H_2 + 2\ OH^-$) rather than Na^+.
  - If $CuSO_4$(aq) with Cu electrodes: $Cu^{2+} + 2\ e^- \rightarrow$ Cu(s) plates on cathode.
- **Anode (oxidation)**:
  - NaCl(aq): often Cl_2 forms ($2\ Cl^- \rightarrow Cl_2 + 2\ e^-$) if [$Cl^-$] high and inert anode; otherwise O_2 from water ($2\ H_2O \rightarrow O_2 + 4\ H^+ + 4\ e^-$).
  - $CuSO_4$(aq): with inert anode, water/OH^- oxidized to O_2; with Cu anode, Cu(s) → $Cu^{2+} + 2\ e^-$ (anode dissolves).

Standard hydrogen electrode (SHE) (HL)

- **SHE half-equation** (reduction form): $2\ H^+(aq) + 2\ e^- \rightleftharpoons H_2(g)$ assigned E° = 0.00 V.
- **Interpreting tables**:
  - more positive E°(reduction) → stronger oxidizing agent (easier to reduce)
  - more negative → stronger reducing agent.

Exam Tip!

Compare two reductions; the more positive runs as reduction, the other is reversed to give oxidation.

Link between ΔG° and E°cell (HL)

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

- where n = moles of electrons transferred; F ≈ 96485 C mol^{-1}.
- **Sign test**: E°cell>0 → ΔG°<0 (spontaneous).
- Convert J→kJ at the end.

Electroplating (HL)

- **Setup**: object to be plated is cathode (−); source metal is often anode (+) (active), solution contains metal ions M^{n+}.
- **Half-equations** (example: Cu plating with Cu anode):
  - **Cathode**: $Cu^{2+} + 2\ e^- \rightarrow$ Cu(s) (coating the object)
  - **Anode** (active): Cu(s) → $Cu^{2+} + 2\ e^-$ (replenishes ions)
  - If anode is **inert**, add metal salt periodically to maintain [M^{n+}].

IB CHEMISTRY R3.3 ELECTRON SHARING REACTIONS

Radicals

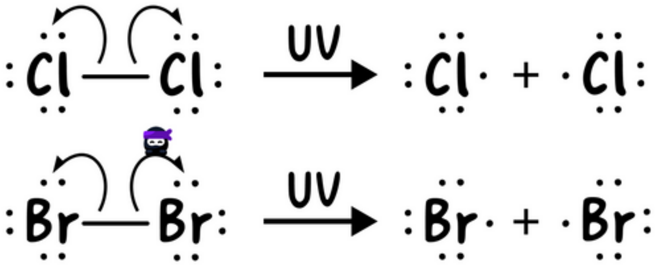
- **Radical** = species with an unpaired electron (•); highly **reactive**; **seeks pairing** via reactions
- **Notation/use**: place a **dot** next to the atom carrying the unpaired electron, e.g., $Cl\bullet$, $CH_3\bullet$, tert-butyl $\bullet C(CH_3)_3$
- **Representation in mechanisms**: use a **single-barbed "fish-hook" arrow** to show movement of one electron

Exam Tip!

Always show the **dot on the correct atom** and conserve total electron count in each step.

Homolytic fission & initiation (UV/heat)

- **Homolytic fission** = breaking a covalent bond so each atom takes one electron; needs **energy** (UV/Δ); generates **two radicals**
- Initiation step for halogens under UV:
  - $Cl_2 \rightarrow (h\nu) \rightarrow 2Cl\bullet$
  - $Br_2 \rightarrow (h\nu) \rightarrow 2Br\bullet$



Radical substitution of alkanes

- Alkanes are unreactive (strong C–H/C–C, non-polar) so they need **radical conditions** (UV/Δ) to undergo substitution with X_2
- **General pattern with methane (chlorination shown)**:
  - Propagation 1: $Cl\bullet + CH_4 \rightarrow HCl + CH_3\bullet$
  - Propagation 2: $CH_3\bullet + Cl_2 \rightarrow CH_3Cl + Cl\bullet$
  - Net(change): $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$
  - Termination (combine any two radicals to remove them):
    - $CH_3\bullet + Cl\bullet \rightarrow CH_3Cl$
    - $CH_3\bullet + CH_3\bullet \rightarrow C_2H_6$
    - $Cl\bullet + Cl\bullet \rightarrow Cl_2$
- **Product mixtures & further substitution:**
  - Once CH_3Cl forms, it can be further chlorinated: $CH_3Cl \rightarrow CH_2Cl_2 \rightarrow CHCl_3 \rightarrow CCl_4$ (under prolonged UV/Cl_2 excess)
  - In higher alkanes, positional isomers form (e.g., chlorination of propane yields 1-chloropropane and 2-chloropropane)
- **Selectivity & stability:**
  - C–H abstraction forms alkyl radicals;
  - **Stability order**: tertiary > secondary > primary > methyl (more substituted radicals are lower in energy)
  - **Br_2** is **more selective** than Cl_2 because endothermic H-abstraction step for $Br\bullet$ has a larger Ea and favors formation at the most stable radical site)
  - **Cl_2** is **less selective/more reactive**

Common Mistake

Don't leave a radical unaccounted for in **propagation** products.

IB CHEMISTRY R3.4 ELECTRON-PAIR SHARING REACTIONS

Nucleophiles

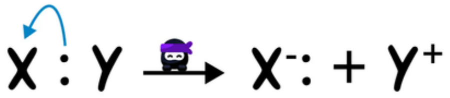
- Definition: A **nucleophile** donates a lone pair to make a new σ bond with an electrophile.
- **Typical nucleophiles (neutral)**: H_2O, NH_3, ROH, RNH_2.
- **Typical nucleophiles (anionic)**: HO^-, RO^-, CN^-, HS^-, $Cl^-/Br^-/I^-$, NO_2^-, NH_2^-.
- **Recognition cues**: lone pair(s) and/or negative charge; often attack δ^+ carbons (e.g., C–X in alkyl halides, C=O carbon).

Exam Tip!

Quickly justify: "X has a lone pair/negative charge → donates e^- pair → nucleophile."

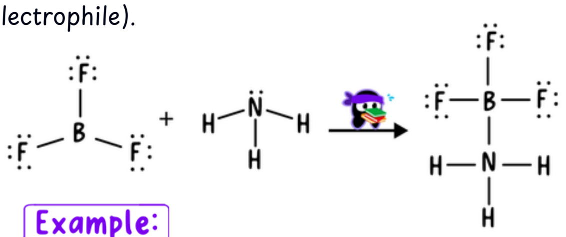
Heterolytic fission

- **Definition**: both bonding electrons go to one fragment when a σ bond breaks → ions.
- **Generic**: $R\text{–}X \rightarrow R^+ + X^-$ (electrons move toward X if X is more electronegative).
- **Generic equations**:
    - $R\text{–}Cl \rightarrow R^+ + Cl^-$ (curly arrow from C–Cl bond to Cl).
    - Proton transfer step: $B:^- + H\text{–}A \rightarrow B\text{–}H + A:^-$ (arrow from B lone pair to H; arrow from H–A bond to A).



Lewis acid and base identification (HL)

- **Lewis base** = electron-pair donor (often the nucleophile).
- **Lewis acid** = electron-pair acceptor (often the electrophile).

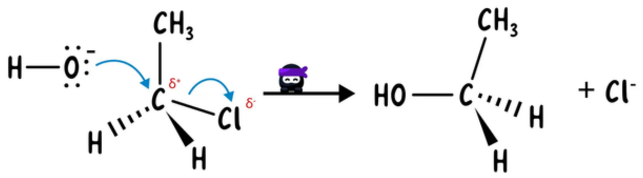


Example:

- **Inorganic** examples:
    - NH_3 (base) + BF_3 (acid) → $F_3B{\leftarrow}NH_3$
    - H_2O (base) + $Al^{3+}(aq)$ (acidic center) → $[Al(H_2O)_6]^{3+}$
- **Organic** examples:
    - Carbonyl O donates to BCl_3 (acid); amines donate to metal centers.

Nucleophilic substitution

- **Core idea**: Nu donates an electron pair to $C\delta^+$ as C–LG bond breaks → leaving group (LG) departs with the bonding pair.
- **Generic equation**: $Nu:^- + R\text{–}LG \rightarrow Nu\text{–}R + LG^-$ (heterolytic C–LG cleavage).

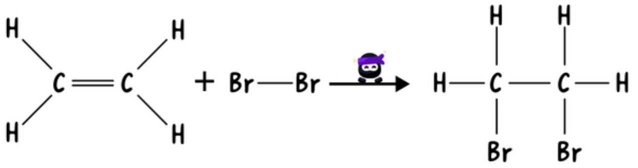


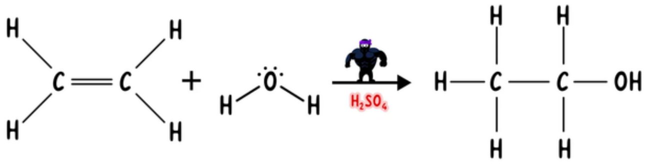
Electrophiles

- **Definition**: An **electrophile** accepts a lone pair from a nucleophile.
- **Typical electrophiles (neutral):** C=O carbon in aldehydes/ketones; SO_3; $BF_3/AlCl_3$ (Lewis acids).
- **Typical electrophiles (cationic/polarized):** H^+/H_3O^+, NO_2^+, carbocations (R_3C^+), polarized C–X ($C\delta^+$).
- **Recognition cues**: positive charge, strong δ^+ center, electron-deficient atoms (incomplete octet).

Electrophilic addition to alkenes

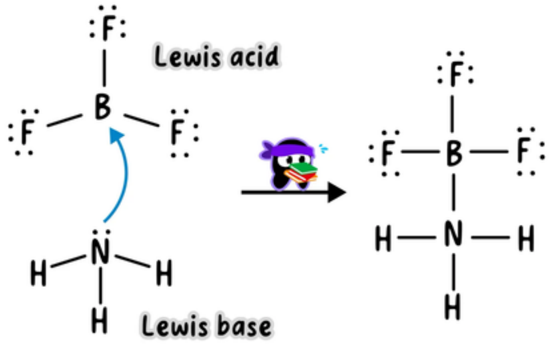
- **Alkenes** react with **electrophiles** at the C=C (π electrons donate).
- **Halogenation**: $C{=}C + X_2 \rightarrow X\text{–}C\text{–}C\text{–}X$ (e.g., ethene + Br_2 → 1,2-dibromoethane; π electrons attack $X\delta^+$, X–X breaks).
- **Hydrohalogenation**: C=C + HX → haloalkane (Markovnikov product in unsym. alkenes).
- **Hydration (overall)**: $C{=}C + H_2O \rightarrow$ alcohol (via H^+ catalysis: first add H^+, then H_2O attacks carbocation, deprotonation gives ROH).





Coordination bond formation (HL)

- When a **Lewis base** donates a lone pair to a **Lewis acid**, a **coordinate (dative) bond** forms.
- Draw as an arrow from donor to acceptor in the product (often later shown as a normal line):
  - $BF_3 + :NH_3 \rightarrow F_3B \leftarrow NH_3$
  - $[Cu(H_2O)_4]^{2+} + 2\ NH_3 \rightarrow [Cu(H_2O)_2(NH_3)_2]^{2+} + 2\ H_2O$ (N donates to Cu^{2+})
- In **proton transfers**, the N→H or O→H donation step is also a coordination bond (e.g., $NH_3 + H^+ \rightarrow NH_4^+$).

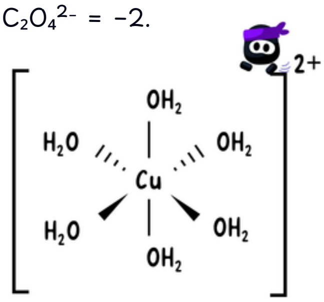


Leaving-group effects on substitution rate (HL)

- **Better leaving group** = weaker base & more stable anion → faster substitution.
- Trend (same R):
  - R–I > R–Br > R–Cl ≫ R–F (I^- best, F^- worst) due to bond strength and anion stability.
- For a given halogen, tertiary > secondary > primary for SN1; primary > secondary > tertiary for SN2 (sterics).

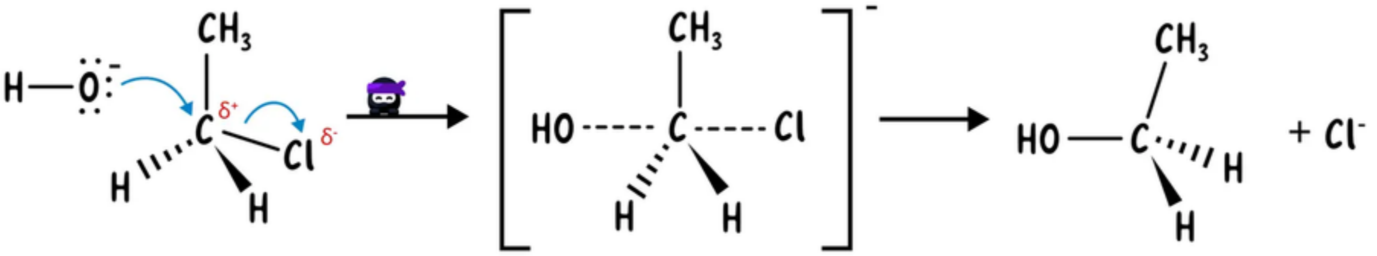
Ligands and coordination bond (HL)

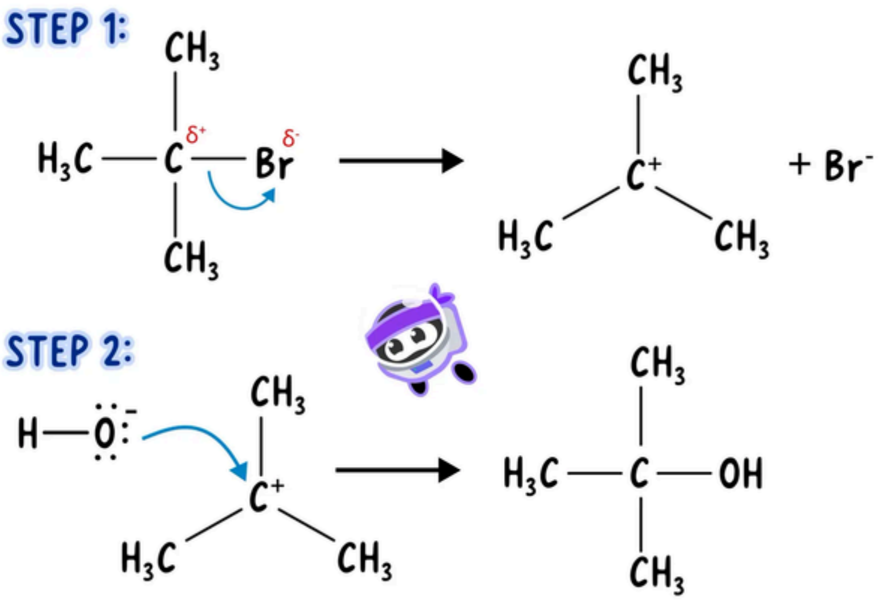
- **Ligands** donate a lone pair to a metal cation → coordinate (dative) bonds
- **Common ligand charges**: H_2O, NH_3, CO, en (ethylenediamine) = 0; Cl^-, Br^-, I^-, CN^-, OH^-, NO_2^- = −1; oxalate $C_2O_4^{2-}$ = −2.



Nucleophilic substitution of halogenoalkanes (HL)

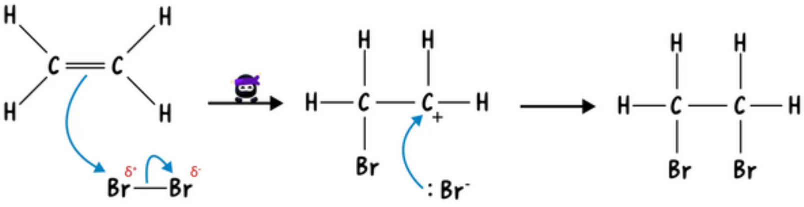
- **SN2 (primary RX)**:
  - bimolecular
  - concerted one-step backside attack
  - transition state with partial bonds
  - inversion of configuration (stereospecific)
  - bimolecular
- **SN1 (tertiary RX)**:
  - unimolecular
  - two-step via planar carbocation
  - nucleophile attacks from either side → racemization (often partial)
- **Secondary halogenoalkanes**: can go SN1 and/or SN2 depending on conditions (not examined here).

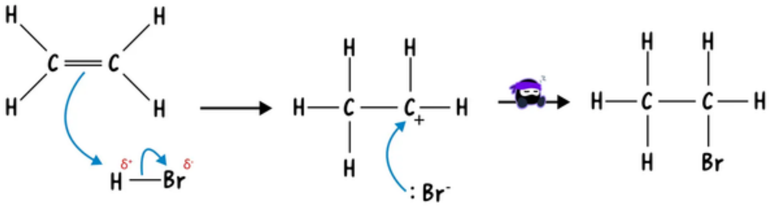


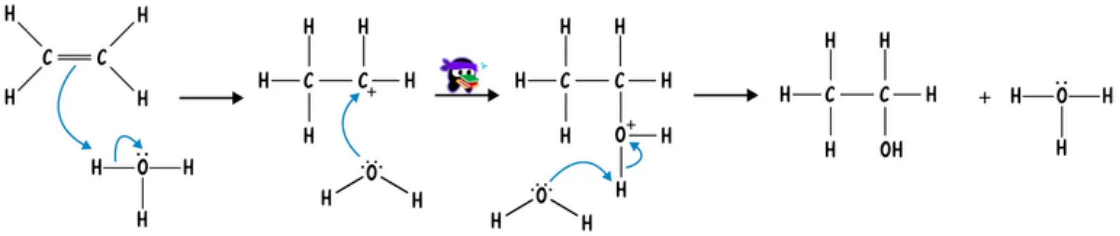


Electrophilic addition(HL)

- **Halogenation**: C=C donates to X_2 → cyclic halonium (not required to name), then X^- opens → vicinal dihalide. E.g., ethene + Br_2 → 1,2-dibromoethane.
- **Hydrohalogenation (HX):** π bond → H^+ adds, then X^- adds → haloalkane; for symmetrical alkenes only one product.
- **Hydration (acid-catalyzed)**: C=C + H_2O (H^+ present) → alcohol.

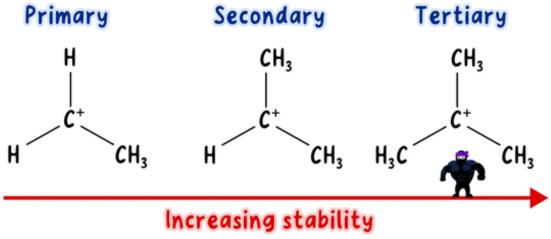






Carbocation stability (HL)

- **Carbocation stability**: 3° > 2° > 1° > methyl (hyperconjugation/inductive effects).
- **Markovnikov orientation**: In HX addition to $R^1CH=CHR^2$, H adds to the carbon with more H's; the more substituted carbocation forms, then X^- attacks → major product.
- **Water addition** (acid-catalyzed) follows analogous logic → Markovnikov alcohol.



Common Question

Which of the following can act as an electrophile?

I. H^+

II. NO_2^+

III. NH_3

A. I and II only

B. I and III only

C. II and III only

D. I, II and III

Electrophilic substitution of benzene (HL)

Steps (generic E^+, formation not examined):

- **Attack**: π system of benzene donates to E^+ → arenium ion (σ-complex; aromaticity lost).
- **Deprotonation**: base removes H^+ from the substituted carbon; aromaticity is restored → substituted benzene + $H\text{-}Base^+$.
- **Arrow pushing**: curved arrow from ring to E^+; then from C–H bond to ring to re-form π system