# Chemistry with Chloe

*Expert A Level Chemistry Tuition*

# AQA A Level Chemistry

## Paper 1: Physical and Inorganic Chemistry

7405/1 • 2 hours • 105 marks • 35% of A-level

## COMPLETE REVISION GUIDE

Physical Chemistry (3.1) • Inorganic Chemistry (3.2)

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# How to Use This Guide

This revision guide covers every specification point examined in AQA A Level Chemistry Paper 1 (7405/1), including sections 3.1.1–3.1.4, 3.1.6–3.1.8, 3.1.10–3.1.12 (Physical Chemistry) and section 3.2 (Inorganic Chemistry).

Each numbered box shows the spec point in italics, followed by a concise, exam-ready explanation. Key terms and phrases are highlighted in bold – these are the words examiners are looking for.

Sections marked with ★ A-LEVEL ONLY are not examined at AS level but are required for the full A-level Paper 1.

Required Practicals are noted where they appear in the specification. Ensure you can describe methods, identify sources of error, calculate uncertainties and suggest improvements.

Use this guide alongside past papers. Good luck! – Chloe

## Key Constants and Standard Values

| Constant / Value | Symbol / Value | Notes |
|---|---|---|
| **Avogadro constant** | NA = 6.022 × $10^{23}$ $mol^{-1}$ | Given in AQA exams |
| **Gas constant** | R = 8.314 J $K^{-1}$ $mol^{-1}$ | Given in AQA exams |
| **Specific heat capacity (water)** | c = 4.18 J $g^{-1}$ $K^{-1}$ | Given if needed |
| **Ionic product of water (25°C)** | Kw = 1.00 × $10^{-14}$ $mol^2$ $dm^{-6}$ | Kw = $[H^+][OH^-]$ |
| **Faraday constant** | F = 96,485 C $mol^{-1}$ | Charge per mole of electrons |
| **Standard conditions** | 100 kPa, 298 K | For all standard enthalpy changes |
| **Standard electrode (SHE)** | E⊖ = 0.00 V | Reference for all E⊖ values |
| **Molar gas volume (RTP)** | 24.0 $dm^3$ $mol^{-1}$ | At 298 K and 100 kPa |

## 3.1.1 Atomic Structure Quick Notes

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **Ar = Σ(mass × abundance) / Σ(abundances)** | Relative atomic mass from mass spectrum data |
| **Protons = Z  Neutrons = A – Z** | Electrons = Z for neutral atom; adjust for ions |
| **Max electrons per shell = $2n^2$** | n = shell number. Shell 1: 2, Shell 2: 8, Shell 3: 18 |
| **KE = ½$mv^2$** | TOF mass spec: all ions given same KE; lighter ions travel faster |

### Electron Configuration

- Fill order: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^{10}$ $4p^6$
- Anomalies: Cr = [Ar] $3d^5$ $4s^1$  Cu = [Ar] $3d^{10}$ $4s^1$
- For d-block IONS: remove 4s electrons FIRST (4s is higher energy in ions)

### Ionisation Energy Trends

- Across Period 3 (Na→Ar): general increase. Drops at Al (3p vs 3s) and S (paired 3p repulsion)
- Down Group 2 (Be→Ba): decrease (more shells, greater shielding, electron further from nucleus)
- Successive IEs: big jump = electron removed from a new inner shell = evidence for shells

## 3.1.2 Amount of Substance Quick Notes

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **n = m / M** | Moles = mass (g) / molar mass (g $mol^{-1}$) |
| **n = c × V** | Moles = concentration (mol $dm^{-3}$) × volume ($dm^3$) |
| **n = V / 24,000  or  n = V / 24** | Gas moles: volume in $cm^3$ ÷ 24,000 or $dm^3$ ÷ 24 at RTP |
| **No. of particles = n × NA** | Number of particles = moles × Avogadro constant |
| **pV = nRT** | Ideal gas: p (Pa), V ($m^3$), n (mol), R (8.314), T (K) |
| **Conc (g $dm^{-3}$) = c (mol $dm^{-3}$) × M** | Convert between concentration units |

### Unit Conversions for pV = nRT

- kPa × 1000 = Pa  $cm^3$ ÷ $10^6$ = $m^3$  $dm^3$ ÷ 1000 = $m^3$  °C + 273 = K

### Yields and Atom Economy

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **% yield = (actual / theoretical) × 100** | Always < 100% (losses, side reactions, equilibrium) |
| **Atom economy = (Mr desired product / ΣMr all reactants) × 100%** | AQA uses REACTANTS in denominator (not products) |

### Equipment Uncertainties

- % uncertainty = (uncertainty / reading) × 100%
- Burette: ±0.05 cm³ per reading (±0.10 cm³ per titre, as two readings are taken)
- 25 cm³ pipette: ±0.06 cm³ 250 cm³ volumetric flask: ±0.10 cm³ Balance: ±0.001 g

## 3.1.3 Bonding Quick Notes

### Shapes of Molecules and Ions (VSEPR)

- 2 electron pairs: linear, 180° ($BeCl_2$, $CO_2$)
- 3 electron pairs: trigonal planar, 120° ($BF_3$). With 1 lp: bent, ~118°
- 4 electron pairs: tetrahedral, 109.5° ($CH_4$, $NH_4^+$). With 1 lp: pyramidal, 107° ($NH_3$). With 2 lp: bent, 104.5° ($H_2O$)
- 5 electron pairs: trigonal bipyramidal, 90°/120° ($PCl_5$)
- 6 electron pairs: octahedral, 90° ($SF_6$)
- Lone pairs reduce bond angles by ~2.5° each. Repulsion: lp–lp > lp–bp > bp–bp

### Intermolecular Forces (weakest to strongest)

- London forces < permanent dipole–dipole < hydrogen bonding
- London forces: present in ALL molecules. Increase with more electrons (higher Mr)
- H-bonding: requires H bonded to N, O, or F + lone pair on N/O/F of adjacent molecule
- Water anomalies: high b.p. (H-bonds); ice less dense than water (open H-bonded lattice)

### Structure and Properties Quick Reference

- High m.p. + conducts when molten/dissolved → GIANT IONIC
- Very high m.p. + hard + doesn't conduct (unless graphite) → GIANT COVALENT
- High m.p. + conducts when solid + malleable → METALLIC
- Low m.p./b.p. + doesn't conduct → SIMPLE MOLECULAR

## 3.1.4 Energetics Quick Notes

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **q = mcΔT** | Energy (J) = mass (g) × c (J $g^{-1}$ $K^{-1}$) × temp change |
| **ΔH = –q / n** | Enthalpy change (kJ $mol^{-1}$). – sign for direction |
| **ΔrH = ΣΔfH(products) – ΣΔfH(reactants)** | Hess's Law using enthalpies of formation |
| **ΔrH = ΣΔcH(reactants) – ΣΔcH(products)** | Hess's Law using enthalpies of combustion |
| **ΔH = Σ(bonds broken) – Σ(bonds formed)** | Mean bond enthalpies (gaseous species only, approximate) |

- ΔfH⊖ of an element in its standard state = 0 by definition
- ΔcH⊖ is always negative (exothermic). ΔfH⊖ can be + or –
- Mean bond enthalpy values give APPROXIMATE ΔH – they are averages across different molecules

## 3.1.6 Chemical Equilibria and Kc Quick Notes

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **$K_c = [C]^c[D]^d / [A]^a[B]^b$** | For aA + bB ⇌ cC + dD. Products over reactants. |

- Le Chatelier's: system shifts to OPPOSE the change
- Increase T → favours endothermic direction. Increase P → favours fewer moles of gas
- Catalyst: reaches equilibrium FASTER but does NOT change position or Kc
- Only TEMPERATURE changes the value of Kc
- Exothermic forward: increase T → Kc decreases. Endothermic forward: increase T → Kc increases
- Units of Kc: substitute (mol $dm^{-3}$) raised to the appropriate powers

## 3.1.7 Redox Quick Notes

- Oxidation = Loss of electrons = Increase in oxidation state (OIL RIG)
- Reduction = Gain of electrons = Decrease in oxidation state
- Oxidation state rules: elements = 0; O = −2 (peroxides = −1); H = +1 (metal hydrides = −1); F = −1 always
- Sum of oxidation states: in a compound = 0; in an ion = charge of the ion
- Half-equations: balance atoms, then balance charges with electrons. Combine so electrons cancel

## ★ 3.1.8 Thermodynamics (A-level only) Quick Notes

### Born–Haber Cycle

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **$\Delta_f H = \Delta_{at}H + IE + EA + \Delta LE$** | Rearrange for the unknown step. Watch signs carefully. |
| **$\Delta_{sol}H = -\Delta LE(\text{dissociation}) + \Sigma\Delta_{hyd}H$** | Enthalpy of solution from lattice + hydration enthalpies |

- AQA: lattice enthalpy can be defined as dissociation (endothermic, +) or formation (exothermic, –). Read the question.
- Experimental LE more exothermic than theoretical → compound has covalent character (polarisation)
- Higher charge density of cation (small, high charge) → more polarising → more covalent character

### Entropy and Gibbs Free Energy

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **$\Delta S_{system} = \Sigma S^{\ominus}(\text{products}) - \Sigma S^{\ominus}(\text{reactants})$** | Units: J $K^{-1}$ $mol^{-1}$. +ΔS = more disorder |
| **$\Delta G = \Delta H - T\Delta S$** | Feasible when $\Delta G \leq 0$ (negative or zero) |
| **$T = \Delta H / \Delta S$** | Temperature at which $\Delta G = 0$ (boundary of feasibility) |

- UNITS: ΔH in kJ $mol^{-1}$ and ΔS in J $K^{-1}$ $mol^{-1}$ – MUST convert ΔS to kJ $K^{-1}$ $mol^{-1}$ (÷ 1000) before using in $\Delta G = \Delta H - T\Delta S$
- Entropy increases: solid → liquid → gas; more moles of gas in products; dissolving
- $\Delta G < 0$ does NOT guarantee the reaction occurs – kinetic barriers (high Ea) may prevent it
- From a graph of ΔG vs T: gradient = –ΔS, y-intercept = ΔH

## ★ 3.1.10 Equilibrium Constant Kp (A-level only) Quick Notes

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **Mole fraction = moles of A / total moles** | Must calculate moles at equilibrium first |
| **Partial pressure = mole fraction × total pressure** | p(A) = x(A) × P(total) |
| **Kp = $p(C)^c \times p(D)^d / p(A)^a \times p(B)^b$** | Partial pressures (atm or Pa). Gases only. |

- Only TEMPERATURE changes Kp. Pressure changes shift the position but NOT Kp
- Catalyst does not affect Kp or position – only speeds up attainment of equilibrium
- Units: substitute pressure units raised to appropriate powers. If powers cancel, Kp has no units

## ★ 3.1.11 Electrode Potentials (A-level only) Quick Notes

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **EMF = E⊖(cathode) – E⊖(anode)** | = E⊖(more positive) – E⊖(less positive) |
| **Positive EMF → reaction is feasible** | Thermodynamically feasible under standard conditions |

### Cell Notation

- Anode (oxidation) on LEFT || Cathode (reduction) on RIGHT (R,O,O,R)
- $Zn(s) | Zn^{2+}(aq) || Cu^{2+}(aq) | Cu(s)$ | = phase boundary || = salt bridge

### Commercial Cells

- Lithium ion: Positive: $Li^+ + CoO_2 + e^- \rightarrow Li^+[CoO_2]^-$ Negative: $Li \rightarrow Li^+ + e^-$
- Alkaline $H_2/O_2$ fuel cell:
- Anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$ Cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
- Overall: $2H_2 + O_2 \rightarrow 2H_2O$ (only product is water)
- Limitations of E⊖: predicts feasibility only (not rate). Non-standard conditions change E

## ★ 3.1.12 Acids and Bases (A-level only) Quick Notes

| Formula / Equation | What It Means / When to Use It |
|---|---|
| **$pH = -\log_{10}[H^+]$** | Definition of pH. Give to 2 d.p. |
| **$[H^+] = 10^{-pH}$** | Convert pH back to $[H^+]$ |
| **Strong acid: $[H^+]$ = conc of acid** | Fully dissociated. pH = –log(conc) |
| **$K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$** | At 25°C. Increases with temperature (endothermic) |
| **Strong base: $[H^+] = K_w / [OH^-]$** | Then pH = –log$[H^+]$ |
| **$K_a = [H^+][A^-] / [HA]$** | Acid dissociation constant for weak acid HA |
| **$[H^+] = \sqrt{(K_a \times [HA])}$** | Assumes $[H^+] = [A^-]$ and [HA] ≈ initial conc |
| **pKa = –log Ka Ka = $10^{-pKa}$** | Lower pKa = stronger weak acid |

### Buffer Solutions

| Formula / Equation | What It Means / When to Use It |
|---|---|
| $[H^+] = Ka \times [HA] / [A^-]$ | Buffer pH: then pH = –log[$H^+$] |
| $pH = pKa + \log([A^-] / [HA])$ | Henderson–Hasselbalch equation |

- Buffer = weak acid + its conjugate base (salt). Resists pH change.
- Action: added $H^+$ reacts with $A^-$. Added $OH^-$ reacts with HA.
- At half-equivalence point of weak acid/strong base titration: [HA] = [$A^-$], so pH = pKa

### Indicator Selection from pH Curves

- Strong acid + strong base: phenolphthalein OR methyl orange (sharp change spans both)
- Weak acid + strong base: phenolphthalein ONLY (equivalence pH > 7)
- Strong acid + weak base: methyl orange ONLY (equivalence pH < 7)
- Weak acid + weak base: NO suitable indicator (no sharp change)
- Phenolphthalein: colourless → pink (~8–10). Methyl orange: red → yellow (~3–5)

## 3.2 Inorganic Chemistry – Key Data

### 3.2.1 Period 3 Trends

- Atomic radius: DECREASES across (more protons, same shielding)
- First IE: general INCREASE across. Drops at Al (3p vs 3s) and S (paired 3p)
- Melting points: Na < Mg < Al (metallic, strengthening). Si very high (giant covalent). $P_4$, $S_8$, $Cl_2$, Ar – simple molecular (low, depends on London forces)

### 3.2.2 Group 2

- Reactivity with water INCREASES down group (IE decreases, easier to lose 2 electrons)
- $M + 2H_2O \rightarrow M(OH)_2 + H_2$ (Mg: slow with cold water, vigorous with steam)
- Hydroxide solubility INCREASES down group: $Mg(OH)_2$ sparingly soluble → $Ba(OH)_2$ soluble
- Sulfate solubility DECREASES down group: $MgSO_4$ soluble → $BaSO_4$ insoluble
- Sulfate test: add dilute HCl, then $BaCl_2$(aq). White ppt of $BaSO_4$ = sulfate present
- Uses: $Mg(OH)_2$ antacid; $Ca(OH)_2$ agriculture; $BaSO_4$ barium meal; Mg extracts Ti from $TiCl_4$

### 3.2.3 Group 7 (Halogens)

- Electronegativity DECREASES down group. Boiling point INCREASES (more electrons → stronger London forces)
- Oxidising ability DECREASES: $Cl_2 > Br_2 > I_2$. $Cl_2$ displaces $Br^-$ and $I^-$
- Reducing ability of halides INCREASES: $I^- > Br^- > Cl^-$
- Solid NaX + conc. $H_2SO_4$: NaCl → HCl (white fumes). NaBr → HBr + $Br_2$ (orange). NaI → HI + $I_2$ (purple) + S + $H_2S$
- Halide test: dilute $HNO_3$ + $AgNO_3$. AgCl = white (dissolves dilute $NH_3$). AgBr = cream (dissolves conc. $NH_3$). AgI = yellow (insoluble in $NH_3$)
- $Cl_2 + H_2O \rightleftharpoons HCl + HOCl$ (disproportionation; Cl: 0 → −1 and +1). HOCl kills bacteria
- $Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$ (bleach)

★ *3.2.4 Period 3 Oxides (A-level only)*

- $Na_2O$, MgO = giant ionic (basic). $Al_2O_3$ = giant ionic (amphoteric). $SiO_2$ = giant covalent (weakly acidic)
- $P_4O_{10}$, $SO_2$, $SO_3$ = simple molecular (acidic). Dissolve in water to form $H_3PO_4$, $H_2SO_3$, $H_2SO_4$
- Trend: metal oxides = basic/amphoteric; non-metal oxides = acidic. $Al_2O_3$ is the key amphoteric oxide

# ★ 3.2.5–3.2.6 Transition Metals (A-level only) Quick Notes

## Vanadium Oxidation States

- $V^{5+}$ ($VO_2^+$) = yellow → $V^{4+}$ ($VO^{2+}$) = blue → $V^{3+}$ = green → $V^{2+}$ = violet
- Reduced stepwise with Zn in acid

## Colour: ΔE = hν = hc/λ

- d-orbital splitting by ligands → d-d transitions absorb visible light → complementary colour seen
- Colour changes when: oxidation state, ligand, or co-ordination number changes (ΔE changes)
- $d^0$ or $d^{10}$ = colourless (no d-d transition possible)

## Hydroxide Precipitates (with NaOH and $NH_3$)

- $Fe^{2+}$ → $Fe(OH)_2$ GREEN ppt. Insoluble in excess NaOH. Insoluble in excess $NH_3$.
- $Fe^{3+}$ → $Fe(OH)_3$ BROWN ppt. Insoluble in excess NaOH. Insoluble in excess $NH_3$.
- $Cu^{2+}$ → $Cu(OH)_2$ BLUE ppt. Insoluble in excess NaOH. DISSOLVES in excess $NH_3$ → deep blue $[Cu(NH_3)_4(H_2O)_2]^{2+}$.
- $Al^{3+}$ → $Al(OH)_3$ WHITE ppt. DISSOLVES in excess NaOH (amphoteric) → $[Al(OH)_4]^-$. Insoluble in excess $NH_3$.

## Ligand Exchange and Shapes

- Octahedral (6 ligands, 90°): $[Cu(H_2O)_6]^{2+}$, $[Fe(H_2O)_6]^{2+}$, $[Co(NH_3)_6]^{2+}$
- Tetrahedral (4 large ligands, 109.5°): $[CuCl_4]^{2-}$, $[CoCl_4]^{2-}$
- Square planar (4 ligands, 90°): cisplatin $[PtCl_2(NH_3)_2]$
- Linear (2 ligands, 180°): $[Ag(NH_3)_2]^+$ (Tollens' reagent)
- Chelate effect: bidentate/multidentate replace monodentate → more free particles → +ΔS → more stable

## Redox Titration Mole Ratios

- $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$ Ratio: 1 $MnO_4^-$ : 5 $Fe^{2+}$
- $2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$ Ratio: 2 $MnO_4^-$ : 5 $C_2O_4^{2-}$
- $KMnO_4$ is self-indicating: purple → colourless. End point = first permanent pink.
- Always acidify with $H_2SO_4$ (never HCl or $HNO_3$)

## Catalysis

- $V_2O_5$ (Contact process): $SO_2 + V_2O_5 \rightarrow SO_3 + V_2O_4$, then $V_2O_4 + ½O_2 \rightarrow V_2O_5$ (regenerated)
- Fe (Haber process): $N_2 + 3H_2 \rightleftharpoons 2NH_3$ (heterogeneous – reaction at surface)
- $Fe^{2+}$ catalyses $I^- + S_2O_8^{2-}$ (homogeneous): $Fe^{2+} \rightarrow Fe^{3+} \rightarrow Fe^{2+}$ (regenerated)
- $Mn^{2+}$ autocatalysis: $MnO_4^- + C_2O_4^{2-}$ reaction starts slow, speeds up as $Mn^{2+}$ product catalyses
- Heterogeneous catalysts can be poisoned (impurities block active sites – cost implication)

# 3.1.1 Atomic Structure

| 3.1.1.1 | *Fundamental particles: protons, neutrons and electrons – relative charge and relative mass.* |
|---|---|
| | **Proton:** relative mass = 1, relative charge = +1. Found in the nucleus.<br>**Neutron:** relative mass = 1, relative charge = 0. Found in the nucleus.<br>**Electron:** relative mass = 1/1836 (≈0), relative charge = −1. Found in shells/orbitals around the nucleus.<br>The nucleus is **small, dense and positively charged**. Most of the atom is empty space. |

| 3.1.1.2a | *Mass number (A) and atomic (proton) number (Z). Determine fundamental particles in atoms and ions. Explain the existence of isotopes.* |
|---|---|
| | **Atomic number (Z):** number of protons. Defines the element.<br>**Mass number (A):** total protons + neutrons.<br>**Neutrons = A – Z.** Electrons = Z for a neutral atom; for ions, adjust by the charge.<br>**Isotopes:** atoms of the **same element** (same Z) with **different numbers of neutrons** (different A). Same chemical properties, different physical properties. |

| 3.1.1.2b | *Principles of a simple time of flight (TOF) mass spectrometer: ionisation, acceleration, ion drift, ion detection, data analysis.* |
|---|---|
| | **1. Ionisation:** sample is ionised (e.g. by electron impact or electrospray) to form positive ions.<br>**2. Acceleration:** ions are accelerated by an electric field so all ions have the **same kinetic energy** ($KE = \frac{1}{2}mv^2$).<br>**3. Ion drift (flight tube):** ions travel through a field-free region. Lighter ions travel **faster** and reach the detector first. Time of flight depends on m/z.<br>**4. Detection:** ions hit a detector, producing a signal proportional to abundance.<br>**5. Data analysis:** m/z and abundance data give a **mass spectrum**. Used to determine relative isotopic masses and abundances.<br>**Ar = Σ(isotopic mass × abundance) / Σ(abundances)** |

| 3.1.1.3 | *Electron configurations of atoms and ions up to Z = 36. First and successive ionisation energies.* |
|---|---|
| | **Fill order:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$. Anomalies: Cr = [Ar] $3d^5\ 4s^1$, Cu = [Ar] $3d^{10}\ 4s^1$.<br>For ions: remove from **4s before 3d**.<br>**First ionisation energy:** energy to remove one mole of electrons from one mole of gaseous atoms: $X(g) \rightarrow X^+(g) + e^-$.<br>**Trends in Period 3 (Na–Ar):** general increase across (more protons, same shielding). Drop at Al (3p easier to remove than 3s) and S (paired 3p electron repulsion).<br>**Trends in Group 2 (Be–Ba):** decrease down group (more shells, greater shielding, electron further from nucleus).<br>**Successive IEs:** large jumps indicate electrons being removed from a **new inner shell** – provides evidence for shell structure. |

# 3.1.2 Amount of Substance

| 3.1.2.1 | *Relative atomic mass and relative molecular mass in terms of $^{12}C$. Relative formula mass for ionic compounds.* |
|---|---|
| | **Relative atomic mass (Ar):** weighted mean mass of an atom of an element compared to 1/12 the mass of a $^{12}C$ atom.<br>**Relative molecular mass (Mr):** sum of Ar values for all atoms in a molecular formula (covalent compounds).<br>**Relative formula mass:** sum of Ar values for all atoms in a formula unit (ionic compounds, e.g. NaCl). |

| 3.1.2.2 | *The Avogadro constant and the mole. Concentration of a substance in solution (mol $dm^{-3}$). Calculations using mass, Mr, moles and concentration.* |
|---|---|
| | **Avogadro constant (NA):** $6.022 \times 10^{23}$ $mol^{-1}$ (number of particles per mole). Given in exams.<br>**Key formulae:**<br>**n = m / M** (moles = mass / molar mass)<br>**n = c × V** (moles = concentration × volume in $dm^3$)<br>**Number of particles = n × NA**<br>Always convert $cm^3$ to $dm^3$ by dividing by 1000. |

| 3.1.2.3 | *The ideal gas equation pV = nRT with variables in SI units.* |
|---|---|
| | **pV = nRT**<br>p = pressure (**Pa**), V = volume (**$m^3$**), n = moles, R = 8.314 J $K^{-1}$ $mol^{-1}$ (given), T = temperature (**K**).<br>**Unit conversions:** kPa × 1000 = Pa; $cm^3 \div 10^6 = m^3$; $dm^3 \div 1000 = m^3$; °C + 273 = K.<br>Used to find Mr of a volatile liquid (given mass, T, P, V). |

| 3.1.2.4 | *Empirical and molecular formula. Calculate from composition by mass or percentage composition.* |
|---|---|
| | **Empirical formula:** simplest whole number ratio of atoms.<br>**Molecular formula:** actual number of atoms in one molecule.<br>**Method:** (1) Divide each mass/% by Ar. (2) Divide all by the smallest. (3) Multiply up to whole numbers if needed.<br>**Molecular formula:** Mr ÷ empirical formula mass = multiplier. |

| 3.1.2.5 | *Balanced equations (full and ionic). Atom economy. Calculations: masses, gas volumes, percentage yields, concentrations.* |
|---|---|
| | **Atom economy = (Mr desired product / ΣMr all reactants) × 100%** (AQA uses reactants in the denominator).<br>**Percentage yield = (actual yield / theoretical yield) × 100%**<br>Yield < 100% due to: incomplete reactions, side reactions, losses during purification.<br>High atom economy is economically and environmentally desirable (less waste).<br>**Titration calculations:** use concordant titres → mean titre → moles of known → mole ratio → moles of unknown → concentration. |

## ⭐ REQUIRED PRACTICAL 1: Make Up a Volumetric Solution and Carry Out a Simple Acid–Base Titration

**Method:**

**1.** Weigh accurately a known mass of solid acid (e.g. potassium hydrogen phthalate) using a balance (±0.001 g). Record mass before and after transfer (by difference).
**2.** Transfer the solid to a beaker. Rinse the weighing bottle into the beaker with distilled water to ensure complete transfer.
**3.** Dissolve in distilled water, stir, then transfer quantitatively to a 250 $cm^3$ volumetric flask using a funnel. Rinse the beaker, rod and funnel into the flask several times.
**4.** Make up to the 250 $cm^3$ graduation mark using a dropping pipette for the final additions. Stopper and invert to mix. This is the standard solution.
**5.** Pipette 25.0 $cm^3$ of the unknown NaOH solution into a conical flask. Add a suitable indicator (e.g. phenolphthalein).
**6.** Fill a burette with the standard acid. Record the initial reading. Titrate, adding acid dropwise near the end point, swirling continuously.
**7.** Record the final burette reading. Repeat to obtain at least two concordant titres (within 0.10 $cm^3$).
**8.** Calculate: mean titre → moles of acid → mole ratio → moles of NaOH → concentration of NaOH.

**Sources of Uncertainty:**

- Balance: ±0.001 g
- Volumetric flask (250 $cm^3$): ±0.10 $cm^3$
- Pipette (25 $cm^3$): ±0.06 $cm^3$
- Burette: ±0.05 $cm^3$ per reading (±0.10 $cm^3$ per titre)

**Common Errors:**

- Not all solid transferred to the volumetric flask (systematic – concentration too low)
- Overshooting the end point (random)
- Misreading the meniscus (parallax error)
- Not making up to the graduation mark accurately
- Rinsing the conical flask with NaOH (increases moles – systematic error)
- Note: a wet conical flask rinsed with distilled water does NOT cause an error

**Improvements:**

- Weigh by difference to account for any solid left in the weighing bottle
- Rinse all glassware into the volumetric flask thoroughly
- Use a dropping pipette for the final approach to the graduation mark
- Use a white tile under the conical flask to see the colour change clearly
- Rinse the burette with the acid solution before filling; rinse the pipette with the NaOH solution
- Add acid dropwise near the end point to avoid overshooting
- Repeat and use concordant titres (within 0.10 $cm^3$) to calculate a mean

# 3.1.3 Bonding

| 3.1.3.1 | *Ionic bonding: electrostatic attraction between oppositely charged ions in a lattice. Formulas of compound ions.* |
|---|---|
| | **Ionic bonding:** the **strong electrostatic attraction** between **oppositely charged ions** arranged in a **giant ionic lattice**.<br>Formed by **electron transfer** from a metal to a non-metal.<br>**Compound ions to know:** sulfate $SO_4^{2-}$, hydroxide $OH^-$, nitrate $NO_3^-$, carbonate $CO_3^{2-}$, ammonium $NH_4^+$.<br>Predict ionic charge from position in the Periodic Table. |

| 3.1.3.2 | *Covalent and dative covalent bonds. Single, multiple and co-ordinate bonds.* |
|---|---|
| | **Covalent bond:** a **shared pair of electrons** between two non-metal atoms. The **strong electrostatic attraction** between the shared electrons and the two nuclei.<br>**Multiple bonds:** double = 2 shared pairs; triple = 3 shared pairs.<br>**Dative covalent (co-ordinate) bond:** both electrons in the shared pair come from **one atom**. Represented by an **arrow** from donor to acceptor (e.g. $NH_4^+$, $Al_2Cl_6$). |

| 3.1.3.3 | *Metallic bonding: attraction between delocalised electrons and positive ions in a lattice.* |
|---|---|
| | **Metallic bonding:** the **strong electrostatic attraction** between **positive metal ions (cations)** and a **sea of delocalised electrons**.<br>Strength increases with: more delocalised electrons, smaller ionic radius, higher charge. This explains trends in m.p. and hardness of metals. |

| 3.1.3.4 | *Four types of crystal structure: ionic, metallic, macromolecular (giant covalent), molecular. Examples and properties.* |
|---|---|
| | **Ionic (e.g. NaCl):** high m.p., conducts when molten/dissolved, brittle.<br>**Metallic (e.g. Mg):** high m.p., conducts when solid, malleable/ductile.<br>**Giant covalent (e.g. diamond, graphite, $SiO_2$):** very high m.p., hard (except graphite), diamond doesn't conduct; graphite does (delocalised electrons in layers).<br>**Simple molecular (e.g. $I_2$, ice):** low m.p./b.p., doesn't conduct. Weak intermolecular forces overcome on melting (NOT covalent bonds).<br>Always link properties to the **type of bonding/structure** and what must be overcome. |

| 3.1.3.5 | *Shapes of simple molecules and ions with up to six electron pairs. VSEPR theory and bond angles.* |
|---|---|
| | **VSEPR:** electron pairs (bonding and lone) repel and arrange to be **as far apart as possible**.<br>**Repulsion order:** lone pair–lone pair > lone pair–bond pair > bond pair–bond pair.<br>**2 pairs:** linear, 180° ($BeCl_2$, $CO_2$).<br>**3 pairs:** trigonal planar, 120° ($BF_3$). With 1 lp: bent, ~118°.<br>**4 pairs:** tetrahedral, 109.5° ($CH_4$, $NH_4^+$). With 1 lp: pyramidal, 107° ($NH_3$). With 2 lp: bent, 104.5° ($H_2O$).<br>**5 pairs:** trigonal bipyramidal, 90°/120° ($PCl_5$).<br>**6 pairs:** octahedral, 90° ($SF_6$).<br>Lone pairs reduce bond angles by ~2.5° each. |

| 3.1.3.6 | *Electronegativity and bond polarity. Polar bonds vs polar molecules.* |
|---|---|
| | **Electronegativity:** the power of an atom to attract the bonding pair of electrons in a covalent bond. Difference in electronegativity → **polar bond** (δ+ and δ–). A molecule is **polar** if the dipoles don't cancel. **Symmetric molecules** (e.g. $CO_2$, $CCl_4$) have polar bonds but are **non-polar overall**. |

| 3.1.3.7 | *Intermolecular forces: permanent dipole–dipole, London forces, hydrogen bonding. Effect on melting/boiling points.* |
|---|---|
| | **London forces (induced dipole–dipole):** present in ALL molecules. Strength increases with number of electrons (higher Mr). Temporary fluctuations in electron density induce dipoles in neighbouring molecules.<br>**Permanent dipole–dipole:** between molecules with permanent dipoles.<br>**Hydrogen bonding:** occurs when H is bonded to **N, O, or F** (highly electronegative atoms with lone pairs). Strongest intermolecular force.<br>**Anomalous properties of water:** high b.p., ice is less dense than liquid water (open H-bonded lattice in ice). |

# 3.1.4 Energetics

| 3.1.4.1 | *Enthalpy change (ΔH). Standard conditions (100 kPa, stated T). Define ΔcH⊖ and ΔfH⊖.* |
|---|---|
| | **Enthalpy change (ΔH):** heat energy change at constant pressure.<br>**Exothermic:** ΔH negative (heat released). **Endothermic:** ΔH positive (heat absorbed).<br>**Standard conditions:** 100 kPa and a stated temperature (usually 298 K).<br>**ΔcH⊖:** enthalpy change when **one mole** of a substance is **completely burned** in excess oxygen under standard conditions.<br>**ΔfH⊖:** enthalpy change when **one mole** of a compound is formed from its **elements in their standard states** under standard conditions. ΔfH⊖ of an element in its standard state = 0. |

| 3.1.4.2 | *Calorimetry: q = mcΔT. Calculate molar enthalpy changes from experimental results.* |
|---|---|
| | **q = mcΔT** where q = energy (J), m = mass of solution (g), c = specific heat capacity (given; for water, 4.18 J $g^{-1}$ $K^{-1}$), ΔT = temperature change.<br>**ΔH = –q / n** (kJ $mol^{-1}$). The negative sign accounts for direction (if solution heats up, reaction is exothermic).<br>**Assumptions:** no heat loss; solution has density/heat capacity of water; all solute reacts. |

| 3.1.4.3 | *Hess's Law: calculate enthalpy changes from enthalpies of formation or combustion.* |
|---|---|
| | **Hess's Law:** total enthalpy change is independent of the route taken.<br>**Using ΔfH⊖:** ΔrH⊖ = ΣΔfH⊖(products) – ΣΔfH⊖(reactants)<br>**Using ΔcH⊖:** ΔrH⊖ = ΣΔcH⊖(reactants) – ΣΔcH⊖(products)<br>Draw an enthalpy cycle showing two routes from reactants to products. |

| 3.1.4.4 | *Mean bond enthalpy. Calculate approximate ΔH for gaseous reactions. Explain why values differ from Hess's Law.* |
|---|---|
| | **Mean bond enthalpy:** average energy required to break one mole of a particular type of bond in the gaseous state, averaged across different molecules.<br>**ΔH = Σ(bonds broken) – Σ(bonds formed)**<br>Bonds broken = endothermic (+). Bonds formed = exothermic (–).<br>**Why values differ from Hess's Law:** mean bond enthalpies are **averages** – actual values depend on molecular environment. Only strictly valid for **gaseous** reactions. |

## ⭐ REQUIRED PRACTICAL 2: Measurement of an Enthalpy Change

**Method:**

**1.** Common examples: neutralisation (NaOH + HCl), displacement (Zn + $CuSO_4$), dissolution ($NH_4NO_3$ or KCl in water), combustion of alcohols (spirit burner).
**2.** FOR NEUTRALISATION/DISPLACEMENT/DISSOLUTION: measure a known volume of solution (e.g. 25 $cm^3$ of 1.0 mol $dm^{-3}$ $CuSO_4$) into a polystyrene cup with a lid.
**3.** Record the initial temperature every minute for 2–3 minutes to establish a baseline.
**4.** Add the second reagent (e.g. excess Zn powder) at a known time. Stir continuously and record the temperature every minute for a further 5–6 minutes.
**5.** Plot temperature vs time. Extrapolate both lines back to the time of mixing to find the true ΔT (correcting for heat loss).
**6.** Calculate $q = mc\Delta T$, then $\Delta H = -q/n$ (kJ $mol^{-1}$). Assign correct sign.
**7.** FOR COMBUSTION: weigh a spirit burner, burn the alcohol under a copper calorimeter containing a known mass of water, record the temperature rise, reweigh the burner. Calculate q and ΔH.

**Sources of Uncertainty:**

- Thermometer: ±0.5°C (standard) or ±0.1°C (digital)
- Measuring cylinder: ±0.5 $cm^3$ (or pipette for greater precision)
- Balance: ±0.001 g
- Polystyrene cup loses heat to surroundings

**Common Errors:**

- Heat loss to the surroundings (systematic – ΔT too small, ΔH less exothermic than true value)
- Assuming the specific heat capacity and density of the solution are the same as pure water
- Not using excess of one reagent (incomplete reaction)
- For combustion: significant heat loss to the air, incomplete combustion, evaporation of alcohol
- Heat absorbed by the cup, thermometer and stirrer (not accounted for)

**Improvements:**

- Use a polystyrene cup with a lid to reduce heat loss
- Use a plotting/extrapolation method (temperature vs time graph) to correct for heat loss
- Use a digital thermometer (±0.1°C) for more precise temperature readings
- Use excess of the solid/second reagent to ensure complete reaction of the solution
- Stir continuously for even heat distribution
- For combustion: use a draught shield, minimise distance between burner and calorimeter
- Repeat and take a mean value for ΔH

# 3.1.6 Chemical Equilibria and Kc

| 3.1.6.1 | *Reversible reactions and dynamic equilibrium. Le Chatelier's principle. Compromise conditions in industry.* |
|---|---|
| | **Dynamic equilibrium** (closed system): forward rate = reverse rate; concentrations remain constant.<br>**Le Chatelier's principle:** if a system at equilibrium is disturbed, it shifts to **oppose the change**.<br>**Temperature:** increase favours the **endothermic** direction. **Concentration:** increase in [reactant] shifts right. **Pressure:** increase shifts towards **fewer moles of gas**. **Catalyst:** no effect on position or K – just reaches equilibrium faster.<br>**Industrial compromise:** e.g. Haber process – moderate T (~450°C) for reasonable rate, high P (~200 atm) for yield, iron catalyst. |

| 3.1.6.2 | *Kc for homogeneous systems. Construct expressions, calculate values, predict effect of temperature.* |
|---|---|
| | For aA + bB ⇌ cC + dD: **$K_c = [C]^c[D]^d / [A]^a[B]^b$**<br>Products over reactants, raised to stoichiometric coefficients.<br>**Only temperature changes Kc.** Changes in concentration or a catalyst do NOT change Kc.<br>**Exothermic forward:** increasing T **decreases** Kc. **Endothermic forward:** increasing T **increases** Kc.<br>Use an **ICE table** (Initial, Change, Equilibrium) to find equilibrium concentrations. Calculate units by substituting (mol $dm^{-3}$) raised to appropriate powers. |

# 3.1.7 Oxidation, Reduction and Redox Equations

| 3.1.7 | *Oxidation and reduction in terms of electron transfer. Oxidation states. Half-equations and overall redox equations.* |
|---|---|
| | **Oxidation = loss of electrons = increase in oxidation state.** (OIL RIG)<br>**Reduction = gain of electrons = decrease in oxidation state.**<br>**Oxidation state rules:** uncombined elements = 0; simple ions = charge; O = −2 (except peroxides −1); H = +1 (except metal hydrides −1); F = −1 always. Sum in a compound = 0; in an ion = charge.<br>**Half-equations:** balance atoms, then balance charges using electrons. Combine half-equations so electrons cancel. |

# 3.1.8 Thermodynamics

★ A-LEVEL ONLY

| 3.1.8.1 | *Born–Haber cycles. Lattice enthalpy, enthalpy of atomisation, ionisation energy, electron affinity. Enthalpies of solution and hydration.* |
|---|---|
| | **Lattice enthalpy:** AQA accepts either lattice dissociation (endothermic, + ) or lattice formation (exothermic, – ) definition. Check the question.<br>**Born–Haber cycle** uses Hess's Law: ΔfH = ΔatH(metal) + ΔatH(non-metal) + IE(s) + EA(s) + ΔLE.<br>Rearrange to find the unknown.<br>**Comparing lattice enthalpies:** if experimental (from Born–Haber) is more exothermic than theoretical (ionic model) → **covalent character** (cation polarises anion).<br>**Enthalpies of solution and hydration:**<br>ΔsolH = –ΔLE(lattice dissociation) + ΣΔhydH<br>OR ΔsolH = ΔLE(lattice formation) + ΣΔhydH. Draw an energy cycle. |

| 3.1.8.2 | *Entropy (ΔS) and Gibbs free energy (ΔG). ΔG = ΔH – TΔS. Feasibility and temperature calculations.* |
|---|---|
| | **Entropy (S):** measure of disorder. Units: J $K^{-1}$ $mol^{-1}$.<br>**ΔSsystem = ΣS⊖(products) – ΣS⊖(reactants)**<br>Entropy increases when: solid → liquid → gas; more moles of gas produced; dissolving.<br>**Gibbs free energy: ΔG = ΔH – TΔS**<br>Feasible when **ΔG ≤ 0**.<br>**Temperature for feasibility:** set ΔG = 0, T = ΔH / ΔS.<br>Ensure consistent units (ΔH in kJ, ΔS in kJ $K^{-1}$, or both in J).<br>ΔG < 0 does NOT guarantee the reaction occurs – kinetic barriers may prevent it. |

# 3.1.10 Equilibrium Constant Kp

★ A-LEVEL ONLY

| 3.1.10 | *Kp for homogeneous gas-phase systems. Partial pressures, mole fractions. Effect of temperature and pressure on Kp.* |
|---|---|
| | **$Kp = p(C)^c \times p(D)^d / p(A)^a \times p(B)^b$** (partial pressures, usually in atm or Pa).<br>**Mole fraction** = moles of A / total moles of gas.<br>**Partial pressure** = mole fraction × total pressure.<br>**Only temperature changes Kp** (same rules as Kc). Concentration, pressure and catalysts do NOT change Kp.<br>Calculate units by substituting pressure units raised to appropriate powers. If Σ(product powers) = Σ(reactant powers), Kp has no units. |

# 3.1.11 Electrode Potentials and Electrochemical Cells

★ A-LEVEL ONLY

| 3.1.11.1 | *Electrode potentials, standard hydrogen electrode, electrochemical series. Calculate EMF. Cell notation.* |
|---|---|
| | **Standard electrode potential (E⊖):** measured under standard conditions (298 K, 100 kPa, 1.00 mol $dm^{-3}$) relative to the **standard hydrogen electrode (SHE)** = 0.00 V.<br>**SHE:** Pt electrode, $H_2$ at 100 kPa, 1.00 mol $dm^{-3}$ $H^+(aq)$, 298 K.<br>**EMF = E⊖(cathode) – E⊖(anode)** = E⊖(more positive) – E⊖(less positive).<br>Positive EMF → reaction is **feasible**.<br>**Cell notation:** anode on LEFT \|\| cathode on RIGHT. E.g. $Zn(s)\|Zn^{2+}(aq)\|\|Cu^{2+}(aq)\|Cu(s)$.<br>**Limitations:** E⊖ only predicts thermodynamic feasibility, not rate. Non-standard conditions change E. |

| 3.1.11.2 | *Commercial cells: non-rechargeable, rechargeable, fuel cells. Lithium ion cells. Alkaline $H_2/O_2$ fuel cell.* |
|---|---|
| | **Non-rechargeable:** irreversible reactions; once reactants used up, cell is dead.<br>**Rechargeable:** reversible reactions; external voltage reverses the reaction.<br>**Lithium ion cell:** Positive: $Li^+ + CoO_2 + e^- \rightarrow Li^+[CoO_2]^-$. Negative: $Li \rightarrow Li^+ + e^-$.<br>**Alkaline $H_2/O_2$ fuel cell:**<br>Anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$. Cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$.<br>Overall: $2H_2 + O_2 \rightarrow 2H_2O$. Only product is water.<br>**Benefits:** clean, efficient. **Risks:** $H_2$ storage (flammable), cost, infrastructure. |

## ⭐ REQUIRED PRACTICAL 8: Measuring the EMF of an Electrochemical Cell

**Method:**

1. Set up two half-cells: e.g. Zn(s) in $ZnSO_4(aq)$ and Cu(s) in $CuSO_4(aq)$, each at 1.00 mol $dm^{-3}$.
2. Connect the two half-cells with a salt bridge (e.g. filter paper soaked in saturated $KNO_3$).
3. Connect the metal electrodes to a high-resistance voltmeter using crocodile clips and wires.
4. Record the voltage (EMF) and note the polarity (+/–) to identify the cathode and anode.
5. Repeat with different half-cell combinations (e.g. $Fe/Fe^{2+}$, $Ag/Ag^+$) to build an electrochemical series.
6. Investigate the effect of changing conditions: vary the concentration of one ion, or change the temperature, and record how the EMF changes.
7. Compare experimental EMF values with calculated E⊖cell values from data book E⊖ values.

**Sources of Uncertainty:**

- Voltmeter: typically ±0.01 V
- Concentration of solutions may not be exactly 1.00 mol $dm^{-3}$
- Temperature may differ from 298 K

**Common Errors:**

- Poor electrical contact between salt bridge and solutions (measured voltage too low)
- Metal electrodes not clean (oxide layer increases resistance, lowers EMF)
- Concentrations not exactly 1.00 mol $dm^{-3}$ (non-standard conditions – E differs from E⊖)
- Salt bridge drying out during measurements
- Drawing current from the cell (causes internal resistance – measured EMF too low)

**Improvements:**

- Sand/polish metal electrodes immediately before use to remove oxide layers
- Use freshly prepared solutions made from accurate weighings
- Ensure the salt bridge makes good contact with both solutions
- Use a high-resistance voltmeter to minimise current draw
- Record the temperature and note any deviation from 298 K
- Repeat each measurement and take a mean

# 3.1.12 Acids and Bases

★ A-LEVEL ONLY

| 3.1.12.1 | *Brønsted–Lowry acids (proton donors) and bases (proton acceptors). Acid–base equilibria involve proton transfer.* |
|---|---|
| | **Acid:** proton ($H^+$) donor. **Base:** proton ($H^+$) acceptor.<br>Acid–base reactions = **proton transfer**.<br>**Conjugate pair:** differs by one $H^+$. E.g. $HA/A^-$ and $H_2O/H_3O^+$. |

| 3.1.12.2 | *pH = $-\log_{10}[H^+]$. Calculate pH of strong acids from concentration.* |
|---|---|
| | **$pH = -\log_{10}[H^+]$** and **$[H^+] = 10^{-pH}$**.<br>For a **strong monobasic acid** (e.g. HCl): $[H^+]$ = concentration of acid (fully dissociated). pH = –log(conc). |

| 3.1.12.3 | *Ionic product of water Kw = $[H^+][OH^-]$. Calculate pH of strong bases.* |
|---|---|
| | **$K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$ mol² dm⁻⁶** at 25°C.<br>For a strong base: $[OH^-]$ = concentration of base. **$[H^+] = K_w / [OH^-]$**, then pH = –log$[H^+]$.<br>Kw increases with temperature (forward dissociation is endothermic). |

| 3.1.12.4 | *Weak acids and Ka. pKa = –log Ka. Calculate pH of weak acids.* |
|---|---|
| | **$K_a = [H^+][A^-] / [HA]$** (acid dissociation constant for weak acid HA).<br>**Assumptions:** $[H^+] = [A^-]$; $[HA] \approx$ initial concentration (dissociation is negligible).<br>**$[H^+] = \sqrt{(K_a \times [HA])}$**, then pH = –log$[H^+]$.<br>**pKa = –log Ka**; $K_a = 10^{-pK_a}$. Lower pKa = stronger weak acid. |

| 3.1.12.5 | *pH curves for all combinations of strong/weak acids and bases. Select indicators from pH curves.* |
|---|---|
| | **Strong acid + strong base:** sharp vertical section; equivalence pH = 7.<br>**Weak acid + strong base:** starts higher; gentle rise then sharp; equivalence pH > 7 (conjugate base hydrolysis); buffer region before equivalence.<br>**Strong acid + weak base:** equivalence pH < 7.<br>**Weak acid + weak base:** no sharp change; no suitable indicator.<br>**Indicator choice:** pKa of indicator must fall within the steep vertical section.<br>Phenolphthalein: colourless → pink (~8–10). Methyl orange: red → yellow (~3–5). |

| 3.1.12.6 | *Buffer solutions: definition, action, pH calculation.* |
|---|---|
| | **Buffer:** resists pH change on addition of small amounts of acid or base.<br>**Acidic buffer:** weak acid + its conjugate base (e.g. $CH_3COOH$ + $CH_3COONa$).<br>**Action:** added $H^+$ reacts with $A^-$; added $OH^-$ reacts with HA.<br>**pH calculation:** $[H^+] = K_a \times [HA] / [A^-]$, then pH = –log$[H^+]$.<br>Or: **$pH = pK_a + \log([A^-] / [HA])$** (Henderson–Hasselbalch). |

## ⭐ REQUIRED PRACTICAL 9: Investigate pH Changes During Acid–Base Titrations

**Method:**

1. Set up a pH meter or data logger with a pH probe. Calibrate using buffer solutions (e.g. pH 4.00 and pH 7.00).
2. Pipette 25.0 $cm^3$ of the acid (e.g. weak acid such as ethanoic acid, or strong acid such as HCl) into a conical flask or beaker.
3. Record the initial pH of the acid solution.
4. Fill a burette with the base (e.g. NaOH). Add the base in small increments (e.g. 1 $cm^3$), swirling after each addition. Record the pH after each addition.
5. Near the equivalence point, add in very small increments (e.g. 0.1–0.2 $cm^3$) to capture the steep change accurately.
6. Continue adding base well beyond the equivalence point (e.g. to 50 $cm^3$ total).
7. Plot pH (y-axis) vs volume of base added (x-axis) to produce the titration curve.
8. Identify the equivalence point (steepest part of the curve) and the buffer region (flat region before equivalence, for weak acid + strong base).
9. Repeat for different combinations: weak acid/strong base, strong acid/weak base.

**Sources of Uncertainty:**

- pH meter: ±0.01 pH units (after calibration)
- Burette: ±0.05 $cm^3$ per reading
- Pipette: ±0.06 $cm^3$
- Temperature affects pH readings and Kw

**Common Errors:**

- pH meter not properly calibrated (systematic – all pH values shifted)
- Not swirling/mixing sufficiently after each addition (pH reading not at equilibrium)
- Taking too-large increments near the equivalence point (missing the steep section of the curve)
- $CO_2$ from the air dissolving in basic solutions (lowers pH slightly)
- Temperature not controlled (Kw varies with temperature)

**Improvements:**

- Calibrate the pH meter with at least two buffer solutions immediately before use
- Swirl thoroughly after each addition and wait for the reading to stabilise before recording
- Use very small increments (0.1–0.2 $cm^3$) around the equivalence point
- Record the temperature
- Use a magnetic stirrer for consistent mixing
- Plot the graph with a smooth curve – do not join points with straight lines
- From the curve: identify equivalence pH, select a suitable indicator, identify the buffer region, and read pH at half-equivalence to find pKa

# SECTION 3.2: INORGANIC CHEMISTRY

## 3.2.1 Periodicity

| 3.2.1.1 | *Classification of elements as s, p, d or f block by position in the Periodic Table.* |
|---|---|
| | Elements classified by the **sub-shell** into which the last electron is added:<br>**s-block:** Groups 1 and 2 (+ H, He). **p-block:** Groups 13–18.<br>**d-block:** transition metals (Groups 3–12). **f-block:** lanthanides and actinides. |

| 3.2.1.2 | *Trends in atomic radius, first ionisation energy and melting point of Period 3 elements (Na–Ar).* |
|---|---|
| | **Atomic radius decreases** across Period 3: more protons, same shielding → greater effective nuclear charge pulls electrons closer.<br>**First IE generally increases** across: more protons, same shielding. Drops at Al (3p vs 3s) and S (paired 3p).<br>**Melting points:** Na < Mg < Al (metallic bonding strengthens: more delocalised electrons, smaller ions). Si very high (giant covalent). $P_4$, $S_8$, $Cl_2$, Ar are simple molecular – low m.p. determined by London forces. |

## 3.2.2 Group 2, the Alkaline Earth Metals

| 3.2.2 | *Group 2 (Mg–Ba): trends in atomic radius, IE, m.p. Reactions with water. Solubility of hydroxides and sulfates. Uses.* |
|---|---|
| | **Atomic radius increases** down Group 2 (more shells). **First IE decreases** (outer electron further from nucleus, more shielded).<br>**Reactions with water:** Mg reacts slowly with cold water, vigorously with steam. Ca–Ba react increasingly vigorously with cold water: $M + 2H_2O \rightarrow M(OH)_2 + H_2$.<br>**Hydroxide solubility increases** down the group: $Mg(OH)_2$ sparingly soluble → $Ba(OH)_2$ soluble.<br>**Sulfate solubility decreases** down the group: $MgSO_4$ soluble → $BaSO_4$ insoluble.<br>**Uses:** $Mg(OH)_2$ in antacids; $Ca(OH)_2$ to raise soil pH in agriculture; $BaSO_4$ in barium meals (X-ray contrast).<br>**Mg used in extraction of Ti** from $TiCl_4$ (Kroll process). **$CaO/CaCO_3$ to remove $SO_2$** from flue gases.<br>**Sulfate test:** add **acidified $BaCl_2$** → white ppt of $BaSO_4$ confirms $SO_4^{2-}$. Acidified with HCl to remove $CO_3^{2-}/SO_3^{2-}$ which would also give precipitates. |

# 3.2.3 Group 7(17), the Halogens

| 3.2.3.1 | *Halogen trends: electronegativity, boiling point, oxidising ability, reducing ability of halides. Tests for halide ions.* |
|---|---|

**Electronegativity decreases** down Group 7 (larger atoms, more shielding).
**Boiling point increases** (more electrons → stronger London forces).
**Oxidising ability decreases** down the group (harder to gain an electron). $Cl_2$ displaces $Br^-$ and $I^-$; $Br_2$ displaces $I^-$.
**Reducing ability of halide ions increases** down the group (larger ions lose electrons more easily). Tested by reaction of solid NaX with conc. $H_2SO_4$:

- NaCl: steamy white fumes (HCl only).
- NaBr: HBr formed, then oxidised to $Br_2$ (orange fumes).
- NaI: HI formed, then oxidised to $I_2$ (purple), S (yellow), $H_2S$ (rotten eggs).

**Halide test:** add dilute $HNO_3$, then $AgNO_3$(aq). AgCl = white (soluble in dilute $NH_3$). AgBr = cream (soluble in conc. $NH_3$). AgI = yellow (insoluble in $NH_3$).

| 3.2.3.2 | *Chlorine with water and with NaOH. Disproportionation. Water treatment. Benefits vs risks.* |
|---|---|

**$Cl_2 + H_2O \rightleftharpoons HCl + HOCl$** (disproportionation: Cl goes from 0 to −1 and +1). HOCl kills bacteria – used in water treatment.
**$Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$** (cold dilute NaOH). NaClO = bleach.
**Benefits of chlorination:** kills bacteria, prevents waterborne diseases. **Risks:** toxic, forms potentially harmful by-products. Society judges benefits outweigh risks.

## ⭐ REQUIRED PRACTICAL 4: Test-Tube Reactions to Identify Cations and Anions

**Method:**

1. CATION TESTS:
2. Group 2 cations: carry out flame tests – clean a nichrome wire in conc. HCl, dip in sample, hold in blue Bunsen flame. Ca = orange-red, Sr = red, Ba = green.
3. $NH_4^+$: add NaOH(aq) and warm gently. Test gas with damp red litmus paper – turns blue if $NH_3$ is present.
4. ANION TESTS:
5. Halide ions ($Cl^-$, $Br^-$, $I^-$): add dilute $HNO_3$ (to remove interfering ions), then $AgNO_3$(aq). AgCl = white, AgBr = cream, AgI = yellow. Confirm with ammonia: AgCl dissolves in dilute $NH_3$, AgBr in conc. $NH_3$, AgI is insoluble.
6. $OH^-$: test pH with universal indicator or pH meter. Alkaline pH confirms $OH^-$.
7. $CO_3^{2-}$: add dilute acid (e.g. HCl). Effervescence ($CO_2$ produced) – bubble through limewater, which turns milky.
8. $SO_4^{2-}$: add dilute HCl (to remove $CO_3^{2-}$/$SO_3^{2-}$), then $BaCl_2$(aq). White precipitate of $BaSO_4$ confirms sulfate.

**Sources of Uncertainty:**

- Visual judgement of flame colours and precipitate colours
- Volume of reagents (consistency between tests)
- Subjective end point for colour changes

**Common Errors:**

- Cross-contamination between test tubes (false positives)
- Not acidifying before the $AgNO_3$ test ($CO_3^{2-}$ or $OH^-$ could give precipitates that are not silver halides)
- Not acidifying with HCl before $BaCl_2$ ($BaCO_3$ or $BaSO_3$ precipitates could give a false positive for sulfate)
- Confusing cream AgBr with white AgCl (similar colours)
- Nichrome wire not cleaned properly between flame tests (previous colour contaminates)

**Improvements:**

- Use clean test tubes and a fresh nichrome wire loop for each test
- Clean the wire loop until no colour is observed in the flame before testing each sample
- Acidify with dilute $HNO_3$ before the silver nitrate test (not HCl, which would add $Cl^-$ ions)
- Acidify with dilute HCl before the $BaCl_2$ test
- Use a white tile or background to observe precipitate colours accurately
- Add ammonia solution to silver halide precipitates to confirm identification
- Carry out tests systematically, recording all observations immediately

# 3.2.4 Properties of Period 3 Elements and Their Oxides

★ A-LEVEL ONLY

| 3.2.4 | *Period 3 elements and oxides: reactions with oxygen, melting points, reactions of oxides with water (pH), reactions with acids/bases.* |
|---|---|
| | **Reactions with $O_2$:** $Na_2O$, MgO, $Al_2O_3$, $SiO_2$, $P_4O_{10}$, $SO_2/SO_3$.<br>**Melting points of oxides:** $Na_2O$, MgO, $Al_2O_3$ are giant ionic (high m.p.; MgO highest due to $Mg^{2+}/O^{2-}$ charges). $SiO_2$ is giant covalent (very high m.p.). $P_4O_{10}$, $SO_2$, $SO_3$ are simple molecular (low m.p.).<br>**Reactions with water and pH:**<br>$Na_2O + H_2O \rightarrow 2NaOH$ (pH ~14). $MgO + H_2O \rightarrow Mg(OH)_2$ (pH ~9, sparingly soluble). $Al_2O_3$: insoluble (amphoteric). $SiO_2$: insoluble.<br>$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$ (pH ~1). $SO_2 + H_2O \rightarrow H_2SO_3$ (pH ~2). $SO_3 + H_2O \rightarrow H_2SO_4$ (pH ~1).<br>**Trend:** metal oxides are **basic/amphoteric**, non-metal oxides are **acidic**.<br>**$Al_2O_3$ is amphoteric:** dissolves in both acid and alkali. |

# 3.2.5 Transition Metals

★ A-LEVEL ONLY

| 3.2.5.1 | *Transition metal properties: complex formation, coloured ions, variable oxidation state, catalytic activity. Ligands and co-ordination number.* |
|---|---|
| | **Transition metal:** d-block element that forms at least one stable ion with an **incomplete d sub-shell**. Sc and Zn are NOT transition metals.<br>**Characteristic properties:** variable oxidation states, coloured compounds, catalytic activity, complex ion formation.<br>**Ligand:** molecule/ion with a lone pair that it donates to a metal ion (dative bond). **Complex:** central metal ion surrounded by ligands. **Co-ordination number:** number of dative bonds to the central ion. |

| 3.2.5.2 | *Ligand substitution reactions. Monodentate ($H_2O$, $NH_3$, $Cl^-$), bidentate (en, $C_2O_4^{2-}$), multidentate ($EDTA^{4-}$). Chelate effect. Haemoglobin and CO.* |
|---|---|
| | **$H_2O$ → $NH_3$ exchange:** no change in co-ordination number (both small, uncharged). May be incomplete (e.g. $[Cu(NH_3)_4(H_2O)_2]^{2+}$).<br>**$H_2O$ → $Cl^-$ exchange:** co-ordination number changes (6→4) because $Cl^-$ is larger. Octahedral → tetrahedral.<br>**Chelate effect:** bidentate/multidentate ligands replace monodentate → increase in number of free particles → **positive ΔS** → more stable (more negative ΔG).<br>**Haemoglobin:** $Fe^{2+}$ complex with a multidentate porphyrin ligand. $O_2$ binds reversibly. **CO is toxic** because it binds more strongly than $O_2$, forming carboxyhaemoglobin (irreversible under normal conditions). |

| 3.2.5.3 | *Shapes of complexes: octahedral, tetrahedral, square planar, linear. Isomerism: cis-trans and optical. Cisplatin.* |
|---|---|
| | **Octahedral:** 6 ligands, 90° (e.g. $[Cu(H_2O)_6]^{2+}$). Can show **cis-trans** isomerism (with monodentate ligands) or **optical** isomerism (with bidentate ligands).<br>**Tetrahedral:** 4 large ligands, 109.5° (e.g. $[CoCl_4]^{2-}$).<br>**Square planar:** 4 ligands, 90° (e.g. cisplatin $[PtCl_2(NH_3)_2]$). **Cisplatin** (cis isomer) is used in cancer treatment – binds to DNA, prevents replication.<br>**Linear:** 2 ligands, 180° (e.g. $[Ag(NH_3)_2]^+$ in Tollens' reagent). |

| 3.2.5.4 | *Colour in transition metal complexes. ΔE = hv. Effect of oxidation state, co-ordination number and ligand on colour.* |
|---|---|
| | **Colour arises** from **d-d transitions**: ligands split d-orbitals into two energy levels. Electrons are promoted by absorbing **visible light** of frequency v where **ΔE = hv = hc/λ**. The colour seen is the **complementary colour** of the light absorbed.<br>**Colour changes** when ΔE changes due to: change in **oxidation state**, change in **ligand**, change in **co-ordination number**.<br>Ions with $d^0$ or $d^{10}$ are **colourless** (no d-d transition possible).<br>**Colorimetry:** used to determine concentration of coloured ions (absorbance ∝ concentration). |

| 3.2.5.5 | *Variable oxidation states. Vanadium (V to II). Redox titrations: $Fe^{2+}/MnO_4^-$ and $C_2O_4^{2-}/MnO_4^-$.* |
|---|---|
| | **Vanadium** reduced by Zn in acid: $VO_2^+$ (yellow, +5) → $VO^{2+}$ (blue, +4) → $V^{3+}$ (green, +3) → $V^{2+}$ (violet, +2).<br>**Redox titrations:**<br>$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$ (1:5 ratio). Self-indicating: first permanent pink = end point.<br>$2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$ (2:5 ratio). Warm to ~60°C. $Mn^{2+}$ autocatalyses.<br>Always use $H_2SO_4$ to acidify (never HCl or $HNO_3$). |

| 3.2.5.6 | *Catalysts: heterogeneous and homogeneous. $V_2O_5$ in Contact process, Fe in Haber process, $Fe^{2+}$ catalysis, $Mn^{2+}$ autocatalysis.* |
|---|---|
| | **Heterogeneous catalyst:** different phase from reactants. Reaction at the surface. Increased surface area = better. Can be **poisoned** (impurities block active sites).<br>**$V_2O_5$ (Contact process):** $SO_2 + V_2O_5 \rightarrow SO_3 + V_2O_4$ (V: +5→+4), then $V_2O_4 + ½O_2 \rightarrow V_2O_5$ (regenerated).<br>**Fe (Haber process):** $N_2 + 3H_2 \rightleftharpoons 2NH_3$.<br>**Homogeneous catalyst:** same phase. Works via intermediate.<br>**$Fe^{2+}$ catalyses $I^-/S_2O_8^{2-}$:** Step 1: $2Fe^{2+} + S_2O_8^{2-} \rightarrow 2Fe^{3+} + 2SO_4^{2-}$. Step 2: $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$. $Fe^{2+}$ regenerated.<br>**$Mn^{2+}$ autocatalysis:** product catalyses its own reaction. $MnO_4^-/C_2O_4^{2-}$ starts slow, speeds up. |

# 3.2.6 Reactions of Ions in Aqueous Solution

★ A-LEVEL ONLY

| 3.2.6 | *Reactions of aqua ions: $[M(H_2O)_6]^{2+}$ (Fe, Cu) and $[M(H_2O)_6]^{3+}$ (Al, Fe) with $OH^-$, $NH_3$ and $CO_3^{2-}$. Amphoteric behaviour.* |
|---|---|
| | **Acidity:** $[M(H_2O)_6]^{3+}$ is **more acidic** than $[M(H_2O)_6]^{2+}$ (higher charge/size ratio of $M^{3+}$ polarises O–H bonds more, making $H^+$ release easier).<br>**With NaOH:**<br>$Fe^{2+} \rightarrow Fe(OH)_2$ green ppt (insoluble in excess). $Fe^{3+} \rightarrow Fe(OH)_3$ brown ppt (insoluble). $Cu^{2+} \rightarrow Cu(OH)_2$ blue ppt (insoluble). $Al^{3+} \rightarrow Al(OH)_3$ white ppt (**dissolves in excess NaOH** – amphoteric: $Al(OH)_3 + OH^- \rightarrow [Al(OH)_4]^-$).<br>**With $NH_3$:**<br>Same initial precipitates. $Cu(OH)_2$ **dissolves in excess $NH_3$** $\rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+}$ (deep blue, ligand exchange). $Al(OH)_3$ and Fe hydroxides are **insoluble in excess $NH_3$**.<br>**With $Na_2CO_3$:** all form precipitates + $CO_2$ gas (for 3+ ions the $CO_3^{2-}$ also acts as a base due to the acidity of $[M(H_2O)_6]^{3+}$). |

## ⭐ REQUIRED PRACTICAL 11: Test-Tube Reactions to Identify Transition Metal Ions in Aqueous Solution

**Method:**

1. Prepare separate test tubes containing small volumes (~2 $cm^3$) of aqueous solutions of: $FeSO_4$ ($Fe^{2+}$), $FeCl_3$ ($Fe^{3+}$), $CuSO_4$ ($Cu^{2+}$), and $AlCl_3$ ($Al^{3+}$).
2. TEST WITH NaOH(aq): add NaOH dropwise to each. Record the colour of the precipitate formed. Then add excess NaOH and record whether the precipitate dissolves.
3. $Fe^{2+} \rightarrow$ green ppt $Fe(OH)_2$ (insoluble in excess). $Fe^{3+} \rightarrow$ brown ppt $Fe(OH)_3$ (insoluble in excess). $Cu^{2+} \rightarrow$ blue ppt $Cu(OH)_2$ (insoluble in excess). $Al^{3+} \rightarrow$ white ppt $Al(OH)_3$ (dissolves in excess NaOH – amphoteric).
4. TEST WITH $NH_3$(aq): repeat with dilute ammonia solution. Add dropwise first, then in excess.
5. Same initial precipitates. In excess $NH_3$: $Cu(OH)_2$ dissolves to form deep blue $[Cu(NH_3)_4(H_2O)_2]^{2+}$ (ligand exchange). Fe and Al hydroxides are insoluble in excess $NH_3$.
6. TEST WITH $Na_2CO_3$(aq): add sodium carbonate solution. All form precipitates. With $M^{3+}$ ions, effervescence ($CO_2$) also occurs because $[M(H_2O)_6]^{3+}$ is acidic enough to react with $CO_3^{2-}$.
7. Record all observations in a table: ion tested, reagent, initial observation, observation with excess.

**Sources of Uncertainty:**

- Visual judgement of precipitate colours
- Volume of reagents added (consistency)
- Judgement of when ‘excess’ has been added

**Common Errors:**

- Cross-contamination between test tubes
- Confusing green $Fe(OH)_2$ with green $Cr(OH)_3$ (if $Cr^{3+}$ is also being tested)
- Not adding enough NaOH/$NH_3$ to test for excess (missing the dissolution of $Al(OH)_3$ or $Cu(OH)_2$)
- $Fe^{2+}$ oxidising to $Fe^{3+}$ in air (green precipitate turns brown at the surface)
- Not distinguishing between ‘no change in excess’ and ‘not enough excess added’

**Improvements:**

- Use clean test tubes for every test
- Use freshly prepared solutions (especially $Fe^{2+}$, which oxidises readily in air)
- Add reagent dropwise first to see the initial precipitate clearly, then add a clear excess
- Use a white tile or background for accurate colour observation
- Record observations immediately – some precipitates change colour on standing (e.g. $Fe(OH)_2$ darkens)
- Carry out tests systematically in a table format for easy comparison

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