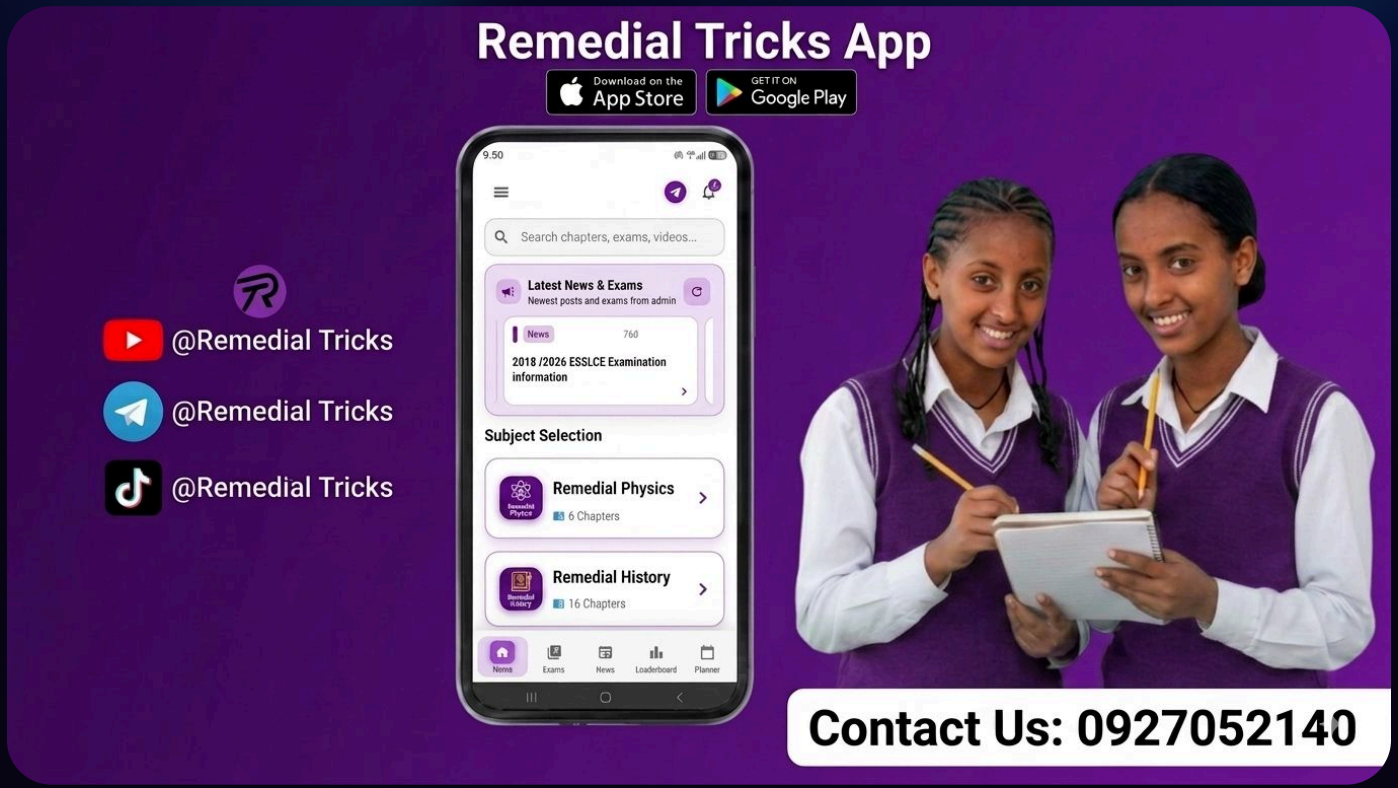
2019 Pre University : Remedial Chemistry course content



⚛️ Unit 1: Atomic Structure and Periodicity

Understand the building blocks of all matter — from subatomic particles to quantum mechanical models and the periodic table's organizing logic.

Topics Covered

- The atom & subatomic particles
- Bohr model of hydrogen
- EM radiation & atomic spectra
- Quantum numbers
- Electronic configuration
- Periodic trends

Key Formula Preview

- $E_n = -13.6 / n^2$ eV
- $c = \lambda\nu$
- $E = h\nu$
- $1/\lambda = R(1/n_1^2 - 1/n_2^2)$

1.1–1.2 The Atom and Subatomic Particles

Core Definitions

Atom

Smallest unit of an element that retains its chemical identity. Composed of a dense nucleus (protons + neutrons) surrounded by electrons.

Element

A pure substance where all atoms have the same atomic number Z. Cannot be broken down by chemical means (e.g., Fe, O, Na).

Compound

Two or more elements chemically bonded in fixed ratios (e.g., H_2O, NaCl). Properties differ from constituent elements.

Subatomic Particles

Particle	Charge	Mass (amu)	Mass (kg)	Location
Proton	+1	1.00728	1.673×10^{-27}	Nucleus
Neutron	0	1.00866	1.675×10^{-27}	Nucleus
Electron	−1	0.000549	9.109×10^{-31}	Orbitals

Atomic Number (Z): Number of protons → uniquely defines the element
Mass Number (A): **A = Z + N** (protons + neutrons)
Neutral atom: electrons = protons = Z

Atomic Mass (Weighted Average)

Atomic mass accounts for the natural abundance of all isotopes:

Atomic mass = Σ (mass × fractional abundance)

Example — Chlorine:
^{35}Cl: 34.969 amu × 0.7577 = 26.496
^{37}Cl: 36.966 amu × 0.2423 = 8.957
Avg atomic mass ≈ 35.45 amu

Isotopes

Isotopes are atoms of the same element (same Z) with different numbers of neutrons (different A). Same chemical behavior, different mass.

Isotope	Symbol	Z	N	Notes
Hydrogen-1	^{1}H	1	0	Protium; 99.98% abundant
Hydrogen-2	^{2}H	1	1	Deuterium; heavy water
Hydrogen-3	^{3}H	1	2	Tritium; radioactive
Carbon-12	^{12}C	6	6	Mass standard (exactly 12 amu)
Carbon-14	^{14}C	6	8	Radiocarbon dating
Uranium-235	^{235}U	92	143	Fissile; nuclear fuel
Uranium-238	^{238}U	92	146	Most abundant U isotope

Historical Atomic Models

01

Dalton (1803)

Solid, indivisible spheres. Elements are atoms; compounds are fixed-ratio combinations.

02

Thomson (1897)

"Plum pudding" — electrons embedded in diffuse positive charge. Discovered via cathode ray tube.

03

Rutherford (1911)

Nuclear model — gold foil experiment revealed small, dense positive nucleus; electrons orbit at distance.

04

Bohr (1913)

Planetary model — electrons occupy fixed energy levels (shells); quantized orbits; explains H spectrum.

Worked Example — ^{56}Fe (Iron-56):
A = 56, Z = 26 → N = A − Z = **30 neutrons**
Neutral atom → electrons = Z = **26 electrons**
Nuclear charge = +26; electron cloud charge = −26

1.3 Bohr Theory of the Hydrogen Atom

Bohr's 4 Postulates

01

Fixed Circular Orbits

Electrons travel in fixed circular orbits around the nucleus without radiating energy. Only certain orbits (n = 1, 2, 3...) are permitted.

02

Quantized Angular Momentum

The angular momentum of an electron is restricted to integer multiples of $\hbar$:

$mvr = n\hbar$, where $\hbar = h/2\pi$ and n = 1, 2, 3...

03

No Radiation in Orbit

An electron in a permitted orbit does **not** radiate energy, despite undergoing centripetal acceleration. Energy is constant in a stationary state.

04

Photon Emission/Absorption

A photon is **emitted** when an electron drops from $n_2 \rightarrow n_1$ ($n_2 > n_1$); a photon is **absorbed** when jumping to a higher level. Energy of photon = ΔE between levels.

Limitations of the Bohr Model

H-like Species Only

Accurate only for one-electron species (H, He^{+}, Li^{2+}). Fails for multi-electron atoms due to electron–electron repulsion.

No Fine Structure

Cannot explain splitting of spectral lines (fine structure). Ignores electron spin and relativistic corrections.

Zeeman Effect

Fails to account for splitting of spectral lines in a magnetic field (Zeeman effect) — requires quantum mechanical treatment.

Key Formulas

Energy of Level n

$E_n = -13.6 / n^2$ eV

n = principal quantum number (1, 2, 3...)

$E_1 = -13.6$ eV (ground state); energy → 0 as n → ∞

Photon Energy

$\Delta E = h\nu = hc/\lambda$

$h = 6.626 \times 10^{-34}$ J·s (Planck's constant)

ν = frequency (Hz); $c = 3.00 \times 10^8$ m/s; λ = wavelength (m)

Rydberg Equation

$1/\lambda = R(1/n_1^2 - 1/n_2^2)$

$R = 1.097 \times 10^7$ m^{-1} (Rydberg constant)

n_1 = lower level; n_2 = upper level ($n_2 > n_1$)

λ = wavelength of emitted/absorbed photon

Hydrogen Spectral Series

Series	n_1	n_2	Region
Lyman	1	2, 3, 4...	Ultraviolet (UV)
Balmer	2	3, 4, 5...	Visible (380–700 nm)
Paschen	3	4, 5, 6...	Infrared (IR)
Brackett	4	5, 6, 7...	Infrared (IR)

Solved Examples

Example 1 — Energy of n = 3 level:

$E_3 = -13.6 / 3^2 = -13.6 / 9 =$ **−1.51 eV**

Example 2 — Balmer transition n = 3 → 2 (wavelength):

$1/\lambda = 1.097 \times 10^7 \times (1/2^2 - 1/3^2)$

$1/\lambda = 1.097 \times 10^7 \times (1/4 - 1/9)$

$1/\lambda = 1.097 \times 10^7 \times (5/36) = 1.524 \times 10^6$ m^{-1}

$\lambda = 1 / 1.524 \times 10^6 =$ **656 nm (red line)**

1.4 Quantum Mechanical Model: EM Radiation & Spectra

Electromagnetic Radiation

Definition

A **transverse wave** consisting of oscillating electric and magnetic fields perpendicular to each other and to the direction of propagation. Travels at **$c = 3 \times 10^8$ m/s** in vacuum.

Wave Properties

λ (wavelength, m): distance between successive crests
ν (frequency, Hz): number of cycles per second
Amplitude: height of wave (determines intensity)
Relationship: $c = \lambda\nu$

Photon Energy

$E = h\nu = hc/\lambda$
$h = 6.626 \times 10^{-34}$ J·s (Planck's constant)
Higher frequency → shorter wavelength → **higher energy**

Wave-Particle Duality

Matter also exhibits wave behavior. **de Broglie wavelength:**
$\lambda = h/mv$
m = mass (kg); v = velocity (m/s)
Significant only for subatomic particles (very small m).

The Electromagnetic Spectrum

Type	Wavelength Range	Energy
Radio	> 1 m	Lowest
Microwave	1 mm – 1 m	Very low
Infrared (IR)	700 nm – 1 mm	Low
Visible	400 – 700 nm	Moderate
Ultraviolet (UV)	10 – 400 nm	High
X-ray	0.01 – 10 nm	Very high
Gamma (γ)	< 0.01 nm	Highest

Atomic Emission Spectra

Continuous spectrum: all wavelengths present (e.g., white light through prism).
Line spectrum (emission): only specific discrete wavelengths emitted — unique to each element, acting as its spectroscopic **"fingerprint"**.

Hydrogen Visible Lines (Balmer Series, $n_1 = 2$)

Colour	λ (nm)	Transition
Violet	410	n = 6 → 2
Blue-green	434	n = 5 → 2
Blue	486	n = 4 → 2
Red	656	n = 3 → 2

Key Formulas

Wave Relationship

$c = \lambda\nu$
$c = 3.00 \times 10^8$ m/s; λ in m; ν in Hz

Photon Energy

$E = h\nu = hc/\lambda$
$h = 6.626 \times 10^{-34}$ J·s

de Broglie Wavelength

$\lambda = h/mv$
m = mass (kg); v = velocity (m/s)

Solved Examples

Example 1 — Photon energy for λ = 486 nm (Balmer blue line):

Step 1: Find frequency
$\nu = c/\lambda = (3.00 \times 10^8) / (486 \times 10^{-9})$
$= 6.17 \times 10^{14}$ Hz

Step 2: Find energy
$E = h\nu = (6.626 \times 10^{-34})(6.17 \times 10^{14})$
$=$ **4.09×10^{-19} J**

Example 2 — de Broglie wavelength of an electron at $v = 2 \times 10^6$ m/s:

Given: $m_e = 9.109 \times 10^{-31}$ kg
$v = 2.00 \times 10^6$ m/s

$\lambda = h / mv$
$\lambda = (6.626 \times 10^{-34}) / (9.109 \times 10^{-31} \times 2.00 \times 10^6)$
$\lambda = (6.626 \times 10^{-34}) / (1.822 \times 10^{-24})$
$=$ **3.64×10^{-10} m = 0.364 nm**

This is on the order of atomic dimensions — confirming that electron wave behavior is physically significant.

Key Insight: The wave-particle duality of electrons underpins the quantum mechanical model, replacing Bohr's fixed orbits with probabilistic orbitals (ψ^2).

1.4.2–1.4.3 Quantum Numbers & Electronic Configuration

The Four Quantum Numbers

Symbol	Name	Meaning	Allowed Values	Example
n	Principal	Energy level; shell size	1, 2, 3, 4…	$n = 3$ → 3rd shell
l	Angular momentum	Subshell shape	0 to $(n-1)$ 0=s, 1=p, 2=d, 3=f	n=3: l = 0,1,2
m_l	Magnetic	Orbital orientation in space	$-l$, …, 0, …, $+l$ ($2l+1$ values)	l=1: m_l = −1,0,+1
m_s	Spin	Intrinsic electron spin	+½ or −½ only	↑ = +½; ↓ = −½

Orbital Shapes & Electron Capacity

l	Subshell	Shape	No. of Orbitals	Max e^-
0	s	Spherical	1	2
1	p	Dumbbell (3 axes)	3	6
2	d	Cloverleaf (4-lobed)	5	10
3	f	Complex multi-lobed	7	14

Three Configuration Rules

Aufbau Principle

Fill orbitals in order of **increasing energy** before higher ones.
Order: 1s → 2s → 2p → 3s → 3p → **4s → 3d** → 4p → 5s → 4d…
Note: 4s fills before 3d due to lower effective energy.

Pauli Exclusion Principle

No two electrons in an atom can have **identical sets of all four quantum numbers**.
Consequence: each orbital holds a **maximum of 2 electrons** with **opposite spins** (+½ and −½).

Hund's Rule

Within a subshell, place **one electron per orbital** before any pairing occurs. All singly occupied orbitals have **parallel spins** (same m_s). Maximises total spin → lower energy state.
e.g., N ($2p^3$): ↑ ↑ ↑ not ↑↓ ↑ _

Periodic Trends

Property	Across Period →	Down Group ↓	Reason
Atomic radius	Decreases ↓	Increases ↑	More protons → stronger nuclear pull; more shells added going down
Ionization energy	Increases ↑	Decreases ↓	Smaller radius → harder to remove e^-; more shielding down group
Electronegativity	Increases ↑	Decreases ↓	Same trends as IE; F is highest (3.98), Fr lowest (~0.7)

Electronic Configuration Examples

Element	Z	Configuration	Period/Group
H	1	$1s^1$	1 / IA
He	2	$1s^2$	1 / VIII A
Li	3	$1s^2\ 2s^1$	2 / IA
C	6	$1s^2\ 2s^2\ 2p^2$	2 / IVA
N	7	$1s^2\ 2s^2\ 2p^3$	2 / VA
O	8	$1s^2\ 2s^2\ 2p^4$	2 / VIA
Na	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$	3 / IA
Cl	17	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	3 / VIIA
Fe	26	$[Ar]\ 3d^6\ 4s^2$	4 / VIII B
Cu ⚠️	29	$[Ar]\ 3d^{10}\ 4s^1$ *(not $3d^9 4s^2$)*	4 / IB

Cu exception: a fully filled $3d^{10}$ subshell is extra stable → one electron migrates from 4s to 3d.

Solved Example — Ca (Z = 20)

Step 1: Apply Aufbau order

1s → 2s → 2p → 3s → 3p → 4s
Count electrons: 2 + 2 + 6 + 2 + 6 + 2 = **20** ✓

Step 2: Write full configuration

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
Condensed: **$[Ar]\ 4s^2$**

Step 3: Identify period & group

Highest n = **4** → **Period 4**
Valence: $4s^2$ (2 valence e^-) → **Group IIA**
Classification: **alkaline earth metal**

Key Insight: Quantum numbers uniquely define every electron's state. Configuration rules (Aufbau + Pauli + Hund) predict ground-state arrangements, with exceptions (Cu, Cr) arising from extra stability of half-filled or fully filled d subshells.

🔗 Unit 2: Chemical Bonding and Structure

Explore how atoms join together — through electron transfer, sharing, or delocalization — and how the arrangement of electrons determines molecular shape and properties.

Topics Covered

- Types of chemical bonds
- Lewis structures & octet rule
- Bond polarity & electronegativity
- VSEPR theory & molecular geometry
- Valence Bond Theory (VBT)
- Molecular Orbital Theory (MOT)

Key Concepts

- Bond polarity: $\Delta\chi$ threshold values
- Hybridization: sp, sp^2, sp^3
- Bond Order = (bonding e^- – antibonding e^-) / 2

2.1–2.2 Types of Chemical Bonding

Ionic Bonding

Definition

Electrostatic attraction between oppositely charged ions, formed by **complete electron transfer** from metal to nonmetal.

Formation steps: (1) metal loses e^- → cation; (2) nonmetal gains e^- → anion; (3) ions attract into crystal lattice.

Lattice energy: energy released forming 1 mol of ionic solid from gaseous ions — higher charge & smaller radius = stronger lattice (MgO > NaCl).

Properties & Examples

- High melting points (strong lattice)
- Brittle (ion displacement → repulsion)
- Conducts electricity when **molten or dissolved** (free ions), not solid

Examples: NaCl ($\Delta\chi$=2.1), MgO ($\Delta\chi$=2.3, very high MP), $CaCl_2$

Covalent Bonding

Definition & Bond Order

Sharing of electron pairs between nonmetal atoms. Bond strength and length vary with order:

Bond	Length (pm)	Energy (kJ/mol)	Example
C–C	154	347	ethane
C=C	134	614	ethene
C≡C	120	839	ethyne

↑ Bond order → shorter & stronger bond

Polar vs Nonpolar & Examples

- **Nonpolar:** $\Delta\chi$ = 0 — H_2, O_2, N_2 (identical atoms, equal sharing)
- **Polar:** $\Delta\chi$ 0.4–1.7 — H_2O, HCl, NH_3 (unequal sharing → dipole)
- CO_2: polar bonds but **nonpolar molecule** (linear, dipoles cancel)

Metallic Bonding

Definition & Properties

"Sea of delocalized electrons" — valence e^- move freely through a lattice of metal cations. This explains:

- **Conductivity:** free e^- carry charge/heat
- **Malleability & ductility:** layers slide without breaking bonds
- **Luster:** free e^- reflect light

Examples: Fe (strong lattice), Cu (high conductivity), Al (lightweight)

Bond Polarity Guide ($\Delta\chi$)

$\Delta\chi$ Range	Bond Type	Example
= 0	Nonpolar covalent	Cl_2
0.01–0.39	Slightly polar covalent	H_2 (~0)
0.4–1.7	Polar covalent	HCl ($\Delta\chi$=0.9)
> 1.7	Ionic	NaCl ($\Delta\chi$=2.1)

Lewis Dot Structures

Rules for drawing:

1. Count total valence e^- (sum all atoms; add 1 per negative charge)
2. Place least electronegative atom in center
3. Connect atoms with single bonds (2e^- each)
4. Distribute remaining e^- as lone pairs — complete octets of outer atoms first
5. If central atom lacks octet, form double/triple bonds
6. Check: total e^- used = total valence e^-

Molecule	Valence e^-	Structure summary
H_2O	8	O center, 2 O–H bonds, 2 lone pairs on O
CO_2	16	C center, 2 C=O double bonds, 2 lone pairs each O
NH_3	8	N center, 3 N–H bonds, 1 lone pair on N

Formal Charge

Used to identify the best Lewis structure (minimize FC magnitudes; negative FC on most electronegative atom).

FC = Valence e^- – Lone pair e^- – ½ Bonding e^-
Sum of all formal charges = overall charge of molecule/ion.

Example (CO_2): C: 4 – 0 – ½(8) = 0 ✓ | O: 6 – 4 – ½(4) = 0 ✓ → FC = 0 on all atoms, confirming the double-bond structure.

2.3–2.4 VSEPR Theory, VBT & MOT

VSEPR — Valence Shell Electron Pair Repulsion

Electron pairs (bonding and lone) repel each other and adopt the geometry that **minimizes repulsion**. Lone pairs occupy more space than bonding pairs.

Repulsion Order

LP–LP > LP–BP > BP–BP
Each lone pair compresses bond angles below ideal.

BP / LP	Geometry	Angle	Example
2 BP / 0 LP	Linear	180°	CO_2, $BeCl_2$
3 BP / 0 LP	Trigonal planar	120°	BF_3
4 BP / 0 LP	Tetrahedral	109.5°	CH_4
3 BP / 1 LP	Trigonal pyramidal	107°	NH_3
2 BP / 2 LP	Bent	104.5°	H_2O
5 BP / 0 LP	Trigonal bipyramidal	90°/120°	PCl_5
6 BP / 0 LP	Octahedral	90°	SF_6

VBT — Valence Bond Theory

Bonds form by **overlap of atomic orbitals**. Two overlap types:

- **σ (sigma):** head-on overlap — all single bonds; one σ per bond
- **π (pi):** side-on overlap — present in double/triple bonds alongside σ

Hybrid	Geometry	Angle	Example
sp	Linear	180°	$BeCl_2$
sp^2	Trigonal planar	120°	BF_3
sp^3	Tetrahedral	109.5°	CH_4
sp^3d	Trig. bipyramidal	90°/120°	PCl_5
sp^3d^2	Octahedral	90°	SF_6

MOT — Molecular Orbital Theory

Atomic orbitals combine to form **molecular orbitals (MOs)** that span the whole molecule:

- **Bonding MOs (σ, π):** lower energy than atomic orbitals → stabilizing
- **Antibonding MOs (σ*, π*):** higher energy → destabilizing

Bond Order = (bonding e^- – antibonding e^-) / 2
Higher bond order = stronger & shorter bond. BO = 0 → molecule does not exist.

Solved Example — O_2 by MOT

MO filling: $\sigma1s^2\ \sigma^*1s^2\ \sigma2s^2\ \sigma^*2s^2\ \sigma2p^2\ \pi2p^4\ \pi^*2p^2$
Bond order = (8 – 4) / 2 = **2** → double bond ✓
O_2 is **paramagnetic** — 2 unpaired electrons in degenerate π* orbitals (confirmed experimentally by attraction to magnets).

🌡️ Unit 3: Physical States of Matter

From ideal gases to crystalline solids — understand how matter behaves at different temperatures and pressures, and the forces that hold it together.

Topics Covered

- Kinetic Molecular Theory
- Gas laws (Boyle, Charles, Gay-Lussac)
- Ideal Gas Law (PV = nRT)
- Dalton's Law of partial pressures
- Surface tension & viscosity
- Intermolecular forces
- Solid types & phase transitions

Core Equation

PV = nRT

R = 0.0821 L·atm/mol·K

This single equation connects pressure, volume, temperature, and moles of any ideal gas.

3.1–3.4 Kinetic Theory & Gas Laws

Kinetic Molecular Theory — 5 Postulates

01

Gas consists of tiny particles in **constant, random motion**

02

Volume of particles is **negligible** compared to the container

03

No intermolecular forces between gas particles

04

Collisions between particles are perfectly **elastic** (no net energy loss)

05

Average kinetic energy is **directly proportional to absolute temperature** (K)

Gas Laws Summary

Law	Formula	Constant	Example
Boyle's	$P_1V_1 = P_2V_2$	T, n	Compress gas → P ↑
Charles's	$V_1/T_1 = V_2/T_2$	P, n	Heat gas → V ↑
Gay-Lussac's	$P_1/T_1 = P_2/T_2$	V, n	Heat sealed tank → P ↑
Avogadro's	$V_1/n_1 = V_2/n_2$	T, P	Add gas → V ↑
Combined	$P_1V_1/T_1 = P_2V_2/T_2$	n	Multi-variable change
Ideal Gas	**PV = nRT**	—	Standard calculations
Dalton's	$P_{total} = P_1 + P_2 + \ldots$	T, V	Gas mixtures

Ideal Gas Law — Variable Definitions

Var	Units	Meaning
P	atm, Pa, mmHg	Pressure of gas
V	L, m^3	Volume of container
n	mol	Moles of gas
R	0.0821 L·atm/mol·K 8.314 J/mol·K	Universal gas constant
T	K (Kelvin)	Absolute temperature

Dalton's Law & Mole Fraction

$P_{total} = P_1 + P_2 + P_3 + \ldots$ — total pressure equals sum of partial pressures.
Mole fraction: **$\chi A = nA/ntotal = PA/Ptotal$**

Graham's Law of Effusion

$rate_1/rate_2 = \sqrt{(M_2/M_1)}$ — lighter gases effuse faster. H_2 effuses ~4× faster than O_2.

Real Gases — van der Waals Equation

Ideal gas law breaks down at high P, low T. Correction:
$(P + an^2/V^2)(V - nb) = nRT$
a = corrects for intermolecular **attractions**; *b* = corrects for finite particle **volume**

Solved Example 1 — Combined Gas Law

A gas occupies **5.0 L** at **300 K** and **1 atm**.
Find volume at **600 K** and **2 atm**.

Using $P_1V_1/T_1 = P_2V_2/T_2$:

$V_2 = P_1V_1T_2 / (T_1 \times P_2)$

$V_2 = (1 \times 5.0 \times 600) / (300 \times 2)$

$V_2 = 3000 / 600 =$ **5.0 L**

Doubling T doubles volume; doubling P halves it — effects cancel out exactly.

Solved Example 2 — Ideal Gas Law

How many moles of gas are in a **10.0 L** container at **2.00 atm** and **350 K**?

PV = nRT → n = PV / RT

n = (2.00 × 10.0) / (0.0821 × 350)

n = 20.0 / 28.735

n = **0.696 mol**

Always convert temperature to Kelvin before using the ideal gas law.

Quick Reference — Key Relationships

Boyle's Law

$P \propto 1/V$
(T, n constant)

Charles's Law

$V \propto T$
(P, n constant)

KE & Temp

$KE_{avg} = (3/2)RT$
per mole

Ideal vs. Real Gases

Condition	Ideal	Real
Particle volume	Zero	Finite
IMFs	None	Present
High T, low P	✓ Valid	≈ Ideal
Low T, high P	✗ Fails	Use vdW

3.5–3.7 Liquids, Solids & Phase Transitions

Intermolecular Forces — Detailed Comparison

Type	Description	Strength	Examples
London Dispersion	Temporary induced dipoles in all molecules	0.05–40 kJ/mol	Noble gases, nonpolar molecules (CH_4, I_2)
Dipole-Dipole	Permanent dipoles attract opposite poles	5–25 kJ/mol	HCl, SO_2, acetone
Hydrogen Bonding	H bonded to N, O, or F attracts lone pair	10–40 kJ/mol	H_2O, HF, NH_3
Ion-Dipole	Ion interacts with polar molecule; strongest of all	Strongest	NaCl dissolved in H_2O

Stronger IMFs → higher BP, higher surface tension, greater viscosity, higher ΔH_vap.

Liquid Properties

Surface Tension

Inward cohesive force at the liquid surface; molecules pulled inward by neighbors. Water = **72.8 mN/m** (exceptionally high due to H-bonding). Increases with IMF strength.

Viscosity

Resistance to flow. Increases with IMF strength and molecular size/complexity. **Honey > Water > Ethanol.** Decreases with temperature.

Capillary Action

Adhesion (liquid–glass attraction) vs. **Cohesion** (liquid–liquid). Adhesion > Cohesion → concave meniscus (water). Cohesion > Adhesion → convex meniscus (mercury).

Solids Classification

Type	Bonding	Properties	MP Range	Examples
Ionic	Ionic bonds	High MP, brittle, molten/aq conducts	600–3000 °C	NaCl, MgO
Covalent Network	Covalent bonds	Very high MP, very hard, non-conductor	>1000°C	Diamond, SiO_2
Metallic	Metallic bonds (sea of e^-)	Variable MP, conductive, malleable	Variable	Fe, Cu, Al
Molecular	IMFs only	Low MP, soft, non-conductor	<300 °C	Ice, I_2, CO_2(s)

Phase Transitions — All 6 Processes

Transition	Direction	Energy	Energy Term
Melting (fusion)	Solid → Liquid	Endothermic	+ΔH_fus
Freezing	Liquid → Solid	Exothermic	−ΔH_fus
Vaporization	Liquid → Gas	Endothermic	+ΔH_vap
Condensation	Gas → Liquid	Exothermic	−ΔH_vap
Sublimation	Solid → Gas	Endothermic	+ΔH_sub
Deposition	Gas → Solid	Exothermic	−ΔH_sub

Hess's Law: ΔH_sub = ΔH_fus + ΔH_vap. Vaporization requires the most energy (breaking all IMFs).

Phase Diagram — Key Features

Triple Point

Unique T and P where all three phases (solid, liquid, gas) coexist in equilibrium. For H_2O: **0.01°C, 0.006 atm.**

Critical Point

Above this T and P, liquid and gas phases are indistinguishable → **supercritical fluid.** For H_2O: 374°C, 218 atm.

Normal BP & MP

Defined at exactly **1 atm.** BP = T where vapor pressure = 1 atm. MP = solid–liquid boundary at 1 atm. For H_2O: BP = 100°C, MP = 0°C.

Slope of Boundaries

Solid–liquid line slopes **left** for H_2O (unusual — ice less dense than water). Slopes **right** for most substances.

Clausius-Clapeyron — Vapor Pressure & Temperature

$$\ln(P_2/P_1) = -\Delta H_{vap}/R \times (1/T_2 - 1/T_1)$$

Vapor pressure increases exponentially with temperature. Higher IMF strength → lower vapor pressure at given T → higher normal boiling point.

🧪 Unit 4: Solutions

Topics Covered

- Homogeneous vs. heterogeneous mixtures
- Types and properties of solutions
- Solubility & "like dissolves like"
- Effects of T and P on solubility
- Henry's Law

Concentration Units

- Molarity (M) = mol / L
- Molality (m) = mol / kg solvent
- Mole fraction (χ) = n_A / n_total
- Mass percent = (m_solute / m_solution) × 100

4.1–4.4 Solution Basics and Solubility

Core Definitions

Solution

Homogeneous mixture with uniform composition throughout; one phase.

Solute

Substance being dissolved (present in smaller amount).

Solvent

Dissolving medium (present in larger amount).

Aqueous Solution

Water is the solvent; most biologically and chemically relevant solutions.

Types of Solutions

Solute	Solvent	Example
Gas	Gas	Air (N_2, O_2)
Gas	Liquid	CO_2 in water
Liquid	Liquid	Ethanol in water
Solid	Liquid	NaCl in water
Solid	Solid	Brass (Cu + Zn)

Dissolution Process (Solvation)

Solvation involves **ion-dipole interactions** between water molecules and ions. For ionic solids:

- **Hydration energy** — energy released when ions are surrounded by water molecules
- **Lattice energy** — energy required to separate ions from the crystal
- If **hydration energy > lattice energy** → compound dissolves

"Like Dissolves Like"

- **Polar** solvents (water) dissolve polar and ionic solutes (NaCl, HCl)
- **Nonpolar** solvents (hexane) dissolve nonpolar solutes (fats, waxes)
- Oil does **not** dissolve in water — nonpolar in polar = immiscible

Factors Affecting Solubility

Temperature:

• Solubility of most solids ↑ **with T** (endothermic dissolution)

• Solubility of gases ↓ **with T** — warm soda goes flat; thermal pollution depletes dissolved O_2

Pressure (gases only):

Henry's Law: C = kP

Gas solubility is directly proportional to its partial pressure above the solution. Negligible effect on solids and liquids.

Solubility Rules for Ionic Compounds

Ion / Group	Solubility	Exceptions
Nitrates (NO_3^-)	✅ Always	None
Chlorides (Cl^-)	✅ Soluble	Ag^+, Pb^{2+}, Hg_2^{2+}
Sulfates (SO_4^{2-})	✅ Soluble	Ba^{2+}, Pb^{2+}, Ca^{2+}
Carbonates (CO_3^{2-})	❌ Insoluble	Group IA, NH_4^+

Colligative Properties (Introduction)

Properties that depend on the **number of solute particles**, not their identity. The van 't Hoff factor **i** accounts for dissociation.

Property	Formula
Boiling Point Elevation	$\Delta T_b = K_b \cdot m \cdot i$
Freezing Point Depression	$\Delta T_f = K_f \cdot m \cdot i$
Osmotic Pressure	$\pi = MRT$

Where **m** = molality, **M** = molarity, **R** = 0.0821 L·atm/mol·K, **i** = van 't Hoff factor (1 for nonelectrolytes; 2 for NaCl, etc.)

4.5–4.8 Concentration & Solution Stoichiometry

Concentration Units — Formulas & When to Use

Unit	Formula	When to Use
Molarity (M)	mol solute / L solution	Most common; temperature-dependent; lab calculations
Molality (m)	mol solute / kg solvent	Temperature-independent; colligative properties (ΔTb, ΔTf)
Mole Fraction (χ)	n_A / (n_A + n_B)	Dimensionless; gas mixtures, Raoult's Law / vapor pressure
Mass Percent (%)	(m_solute / m_soln) × 100	Industrial solutions; label concentrations; density-based
ppm	mg solute / L solution	Trace contaminants; environmental & water quality analysis

Dilution — $M_1V_1 = M_2V_2$

Principle: Moles of solute are conserved upon dilution — only the volume changes.

$M_1V_1 = M_2V_2$ → moles before = moles after

Solved Example: Dilute 25 mL of 6.0 M HCl to a final volume of 150 mL. Find M_2.

$M_2 = M_1V_1 / V_2 = (6.0\ M \times 0.025\ L) / 0.150\ L =$ **1.0 M HCl**

Concentration Interconversion: M ↔ m

Given **density (d, g/mL)** of a solution with molar mass **MM** (g/mol) and molarity **M**:

$m = M / (d - M \times MM/1000)$

Strategy: assume 1 L solution → mass = d × 1000 g; subtract mass of solute = M × MM; remainder is kg of solvent.

Solved Example — Molarity

Dissolve 5.85 g NaCl (MM = 58.5 g/mol) in water to make 500 mL solution:

moles NaCl = 5.85 / 58.5 = 0.10 mol
M = 0.10 mol / 0.500 L = **0.20 M NaCl**

Solution Stoichiometry — 4-Step Method

1 — Write Balanced Equation

Identify all reactants and products with correct coefficients.
$HCl + NaOH \rightarrow NaCl + H_2O$

2 — Find Moles of Known

Convert volume × molarity → moles of known reactant.
moles HCl = 0.025 L × 0.5 M = **0.0125 mol HCl**

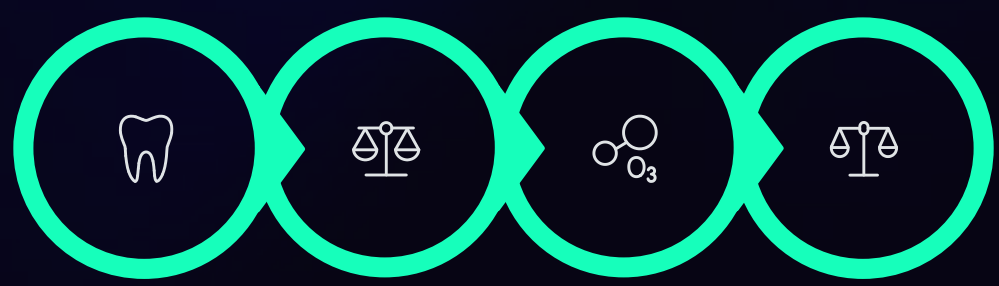
3 — Apply Mole Ratio

Use stoichiometric coefficients from balanced equation.
1 mol HCl : 1 mol NaOH → **0.0125 mol NaOH needed**

4 — Find Volume or Mass of Unknown

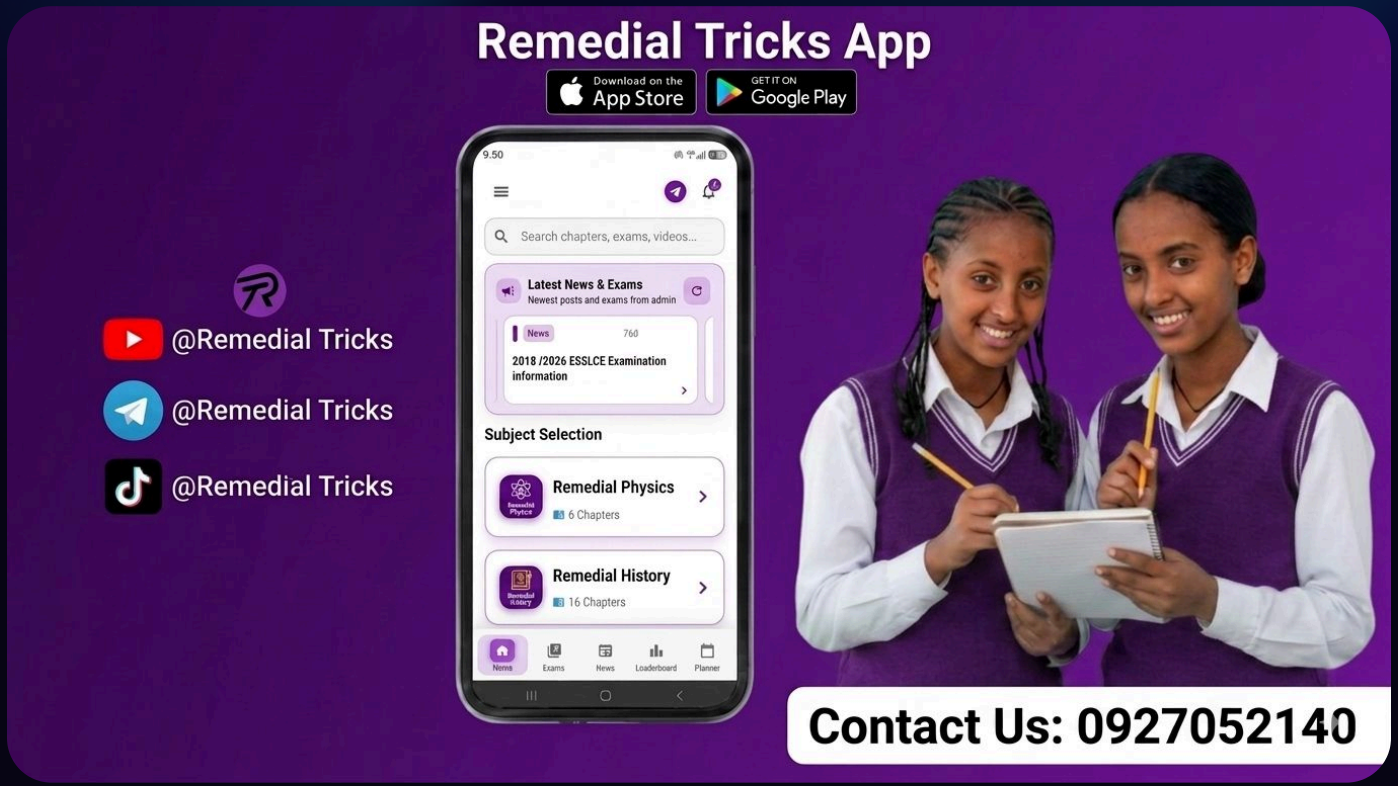
Divide moles by molarity (or multiply by MM for grams).
V = 0.0125 mol / 0.25 M = **50 mL of 0.25 M NaOH**

Stoichiometry Flow (M × V → moles)



Key Rule: Always convert mL → L before computing moles. Molarity uses liters, not milliliters.

FIRST ROUND MODEL
(Units 1–4)



⚗️ Unit 5: Chemical Reactions and Stoichiometry

Balance equations, calculate masses and moles of reactants and products, and understand how energy changes accompany chemical reactions.

Topics Covered

- Chemical equations & balancing
- Laws of conservation, definite & multiple proportions
- Exothermic vs. endothermic reactions
- Hess's Law & enthalpy (ΔH)
- Reaction types (synthesis, decomposition, displacement, combustion)
- Stoichiometric calculations
- Reaction rates

Key Formulas

- $\Delta H_{rxn} = \Sigma \Delta H_f(\text{products}) - \Sigma \Delta H_f(\text{reactants})$
- Moles = mass / molar mass
- Rate = $\Delta[\text{conc.}] / \Delta\text{time}$

5.1–5.4 Equations, Laws & Energy Changes

Chemical Equation Components

Component	Meaning / Example
Reactants	Substances that enter the reaction (left of →)
Products	Substances formed (right of →)
Coefficients	Numbers balancing atom counts (e.g., **4**Fe)
State symbols	(s) solid, (l) liquid, (g) gas, (aq) aqueous

Balancing Equations — Step-by-Step

01

Balance metals first

Adjust coefficients for metallic elements.

02

Then nonmetals

Balance non-hydrogen, non-oxygen elements next.

03

Then hydrogen

Adjust H-containing compounds.

04

Then oxygen

Balance O last — often easiest to fix by fraction then multiply.

Worked Example: $Fe + O_2 \rightarrow Fe_2O_3$
Step 1: Balance Fe → $\mathbf{4}Fe + O_2 \rightarrow \mathbf{2}Fe_2O_3$
Step 2: Balance O → $4Fe + \mathbf{3}O_2 \rightarrow 2Fe_2O_3$ ✓
Check: 4 Fe each side; 6 O each side.

Three Fundamental Laws

Law	Definition	Example
Conservation of Mass (Lavoisier)	Matter is neither created nor destroyed; total mass of reactants = total mass of products.	$2H_2 + O_2 \rightarrow 2H_2O$; mass in = mass out
Definite Proportions (Proust)	A compound always contains elements in fixed mass ratios.	Water is always 11.1% H and 88.9% O by mass.
Multiple Proportions (Dalton)	Elements can combine in different whole-number ratios to form distinct compounds.	CO vs. CO_2 — same elements, different ratios.

Thermochemistry — Core Concepts

Term	Definition
System	The reaction being studied
Surroundings	Everything outside the system
Enthalpy (H)	Heat content of a system at constant pressure
ΔH	**ΔH = H_products − H_reactants**

Exothermic vs. Endothermic

Exothermic (ΔH < 0)	Endothermic (ΔH > 0)
Heat **released** to surroundings	Heat **absorbed** from surroundings
Products more stable (lower energy)	Products less stable (higher energy)
Combustion: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + heat	Photosynthesis: $6CO_2 + 6H_2O$ + light → $C_6H_{12}O_6 + 6O_2$
Neutralization: HCl + NaOH → $NaCl + H_2O$	Dissolving NH_4NO_3 in water (cold pack)

Hess's Law & Standard Enthalpy of Formation

Hess's Law

Enthalpy is a **state function** — ΔH_rxn is the same regardless of pathway taken.
ΔH_rxn = ΣΔH°f(products) − ΣΔH°f(reactants)

Standard Enthalpy of Formation (ΔH°f)

Enthalpy change when **1 mol** of a compound forms from its elements in their standard states.
Rule: ΔH°f of any pure element in standard state = **0**

Solved Example — Combustion of Methane

Step	Calculation
Reaction	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
ΔH°f values (kJ/mol)	CO_2 = −393.5; H_2O = −285.8; CH_4 = −74.8; O_2 = 0
Products sum	[−393.5 + 2(−285.8)] = −965.1 kJ
Reactants sum	[−74.8 + 2(0)] = −74.8 kJ
ΔH_rxn	**−965.1 − (−74.8) = −890.3 kJ/mol** (exothermic)

5.5–5.8 Reaction Types, Balancing & Stoichiometry

Five Reaction Types

Type	Pattern	Balanced Example	Key Indicator
Synthesis	A + B → AB	$2Na + Cl_2 \rightarrow 2NaCl$	Two substances combine into one
Decomposition	AB → A + B	$2H_2O_2 \rightarrow 2H_2O + O_2$	One substance breaks down
Single displacement	A + BC → AC + B	$Zn + 2HCl \rightarrow ZnCl_2 + H_2$	More reactive element displaces less reactive
Double displacement	AB + CD → AD + CB	$AgNO_3 + NaCl \rightarrow AgCl\downarrow + NaNO_3$	Ion exchange; often forms precipitate
Combustion	$C_xH_y + O_2 \rightarrow CO_2 + H_2O$	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$	Hydrocarbon + oxygen

Activity Series (Decreasing Reactivity)

K > Ca > Na > Mg > Al > Zn > Fe > Ni > Sn > Pb > H > Cu > Ag > Au

A metal higher on the list displaces a metal lower on the list from its solution. Metals below H do not react with acids to produce H_2.

Reaction Rate Factors

Factor	Effect
Temperature	↑ temp → ↑ kinetic energy → more frequent, energetic collisions
Concentration	↑ concentration → more particles per volume → more collisions
Surface Area	↑ surface area (smaller particles) → more reactant exposed → faster reaction
Catalyst	Lowers activation energy; not consumed; speeds up reaction without changing products

Stoichiometry — 4-Step Method

01

Balance the equation

Ensure equal atoms on both sides

02

Convert mass → moles

moles = mass ÷ molar mass

03

Apply mole ratio

Use coefficients from balanced equation

04

Convert moles → target unit

mass = moles × molar mass

Solved Example 1 — Simple Stoichiometry

Step	Work
Problem	Grams of H_2O from 4 g H_2 with excess O_2?
Balanced equation	$2H_2 + O_2 \rightarrow 2H_2O$
Moles of H_2	4 g ÷ 2 g/mol = **2 mol H_2**
Mole ratio	2 mol H_2 : 2 mol H_2O → **2 mol H_2O**
Mass of H_2O	**2 × 18 = 36 g H_2O**

Solved Example 2 — Limiting Reagent

Step	Work
Problem	10 g Fe + 8 g S → FeS; find mass of FeS produced
Balanced equation	Fe + S → FeS
Moles of Fe	10 ÷ 55.8 = **0.179 mol**
Moles of S	8 ÷ 32.1 = **0.249 mol**
Limiting reagent	Fe (0.179 mol < 0.249 mol) → **Fe is limiting**
Mass of FeS	**0.179 × 87.9 = 15.7 g FeS**

Percent Yield

Formula	**% yield = (actual yield ÷ theoretical yield) × 100**
Example	Theoretical = 36 g H_2O; Actual = 32 g H_2O
Result	**(32 ÷ 36) × 100 = 88.9%**

⚖️ Unit 6: Chemical Equilibrium

When reactions don't go to completion — dynamic equilibrium, equilibrium constants, and how to predict and control reaction direction using Le Chatelier's Principle.

Topics Covered

- Reversible reactions
- Characteristics of equilibrium
- Equilibrium constant Kc and Kp
- Le Chatelier's Principle
- Industrial applications (Haber Process)

Core Formula

For $aA + bB \rightleftharpoons cC + dD$:

$$K_c = [C]^c[D]^d / [A]^a[B]^b$$

$K > 1 \rightarrow$ Products favored

$K < 1 \rightarrow$ Reactants favored

6.1–6.4 Equilibrium Expressions & Constants

Reversible Reactions & Dynamic Equilibrium

- A **reversible reaction** can proceed in both forward and reverse directions: A ⇌ B
- Equilibrium reached when **forward rate = reverse rate**
- Concentrations remain **constant** but are **not necessarily equal**
- Requires a **closed system**; both reactions continue simultaneously

Equilibrium Constant Kc (Law of Mass Action)

For aA + bB ⇌ cC + dD:

$$K_c = [C]^c[D]^d / [A]^a[B]^b$$

- Only **aqueous** and **gaseous** species included
- Pure solids and pure liquids are **omitted** (activity = 1)

Kp (Partial Pressures)

Uses partial pressures of gases instead of concentrations:

$$K_p = K_c(RT)^{\Delta n}$$

where Δn = moles gas products − moles gas reactants; R = 0.0821 L·atm/mol·K

K Values & Interpretation

K value	Meaning	Direction
K >> 1	Products strongly favored; reaction nearly complete	→ forward
K ≈ 1	Significant amounts of both reactants and products	⇄ balanced
K << 1	Reactants strongly favored; little product forms	← reverse

Reaction Quotient Q

Same expression as Kc but using **non-equilibrium** concentrations:

Condition	Meaning	Shift
Q < K	Too many reactants	→ forward
Q > K	Too many products	← reverse
Q = K	System at equilibrium	⇄ no shift

Kw — Equilibrium of Water

$H_2O \rightleftharpoons H^+ + OH^-$

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

- $pH = -\log[H^+]$
- $pOH = -\log[OH^-]$
- **pH + pOH = 14**

Solved Example 1 — Calculating Kc

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

At equilibrium: $[N_2] = 0.5$ M, $[H_2] = 0.3$ M, $[NH_3] = 0.2$ M

$K_c = [NH_3]^2 / ([N_2][H_2]^3)$

$K_c = (0.2)^2 / (0.5 \times (0.3)^3)$

$K_c = 0.04 / (0.5 \times 0.027)$

$K_c = 0.04 / 0.0135 =$ **2.96**

K > 1 → products (NH_3) slightly favored at these conditions.

Solved Example 2 — Calculating pH

Find the pH of a **0.01 M HCl** solution.

HCl is a strong acid → fully dissociates:

$HCl \rightarrow H^+ + Cl^- \therefore [H^+] = 0.01$ M

$pH = -\log[H^+] = -\log(0.01) = -\log(10^{-2})$

pH = **2.0**

Acidic solution (pH < 7) as expected for a strong acid.

6.5–6.6 Le Chatelier's Principle & Industrial Applications

Le Chatelier's Principle

*"If a system at equilibrium is subjected to a disturbance, the system shifts in the direction that **counteracts the disturbance** and re-establishes equilibrium."*

The system does **not** return to the original equilibrium — it establishes a **new** equilibrium with different concentrations.

Disturbance Effects — Full Table

Disturbance	Shift	Explanation
Add reactant	→ forward	Shift consumes added reactant, forming more product
Remove reactant	← reverse	Shift replenishes lost reactant from products
Add product	← reverse	Shift consumes added product, forming more reactant
Remove product	→ forward	Shift replenishes lost product; drives reaction forward
↑ Pressure	→ fewer gas moles	Reduces volume; shift reduces total moles of gas
↓ Pressure	→ more gas moles	Increases volume; shift increases total moles of gas
↑ Temperature	→ endothermic	Heat acts as reactant (endothermic) or product (exothermic); shift absorbs added heat
↓ Temperature	→ exothermic	Shift releases heat to compensate for heat loss
Add catalyst	⇄ no shift	Speeds both directions equally; reaches equilibrium **faster** but K unchanged

Solved Example — Pressure Shift

Predict the shift when pressure is increased:
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

- Reactant side: 1 + 3 = **4 moles gas**
- Product side: **2 moles gas**
- ↑ Pressure → system shifts toward **fewer moles → forward (→)**
- More NH_3 produced; $Q < K$ until new equilibrium is reached

Industrial Application 1: Haber Process

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92$ kJ/mol

Temperature: 400–500°C — *compromise*: lower T favors NH_3 (exothermic) but reaction too slow; higher T increases rate but decreases yield

Pressure: 150–300 atm — high P favors product side (4 mol gas → 2 mol gas); very high P is costly/dangerous

Catalyst: Iron (Fe) with K_2O/Al_2O_3 promoters — no shift in equilibrium; lowers activation energy for both directions

Importance: Produces NH_3 for nitrogen fertilizers; feeds approximately **50% of the world's population**

NH_3 is continuously **removed** (liquefied) from the reactor — by Le Chatelier's Principle, this drives the reaction **forward**, improving overall yield despite moderate conditions.

Industrial Application 2: Contact Process

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -197$ kJ/mol

Temperature: 450°C — compromise between rate (higher T) and yield (lower T favors exothermic product)

Catalyst: Vanadium pentoxide (V_2O_5) — dramatically increases rate; no effect on K

Pressure: ~1–2 atm — high conversion (~99%) already achieved; high P not economically necessary

Product: SO_3 dissolved in H_2SO_4 → oleum → diluted to produce **sulfuric acid (H_2SO_4)**

Importance: H_2SO_4 is the world's most produced industrial chemical; used in fertilizers, batteries, refining

Key Distinction: K vs. Rate

Factor	Affects K?
Concentration change	No — Q shifts, K unchanged
Pressure change	No — Q shifts, K unchanged
Temperature change	**Yes** — K itself changes
Catalyst	No — only rate changes

🔋 Unit 7: Electrochemistry

Explore the relationship between chemical reactions and electrical energy — from the fundamentals of oxidation-reduction to real-world battery and electroplating applications.

Topics Covered

- Electrolytes and conductivity
- Oxidation and reduction (redox)
- Oxidation numbers
- Voltaic vs. electrolytic cells
- Standard cell potential (E°)
- Faraday's Laws of electrolysis
- Industrial applications

Core Formulas

- $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
- $m = (Q \times M) / (n \times F)$
- $Q = I \times t$
- $F = 96{,}485\ C/mol$

7.1–7.4 Conductivity, Redox & Electrochemical Cells

Electrolytes & Conductivity

Type	Dissociation	Examples
Strong	Fully dissociates	NaCl, HCl, NaOH
Weak	Partially dissociates	CH_3COOH, NH_3
Nonelectrolyte	Does not dissociate	Glucose, ethanol

Conductivity ∝ ion concentration in solution.

Oxidation Number Rules

- Free element = **0** (e.g., Na, Cl_2, O_2)
- Monatomic ion = **its charge** (e.g., Na^+ = +1)
- O = **–2** (except peroxides: –1; OF_2: +2)
- H = **+1** (except metal hydrides: –1)
- Sum of oxidation numbers = **charge of species**

OIL RIG — Redox Fundamentals

OIL — Oxidation Is Loss

Loses e^- → ox. no.
increases → ANODE
e.g., $Zn \rightarrow Zn^{2+} + 2e^-$

RIG — Reduction Is Gain

Gains e^- → ox. no.
decreases → CATHODE
e.g., $Cu^{2+} + 2e^- \rightarrow Cu$

Balancing Redox — Half-Reaction Method

1. Separate into two half-reactions
2. Balance all atoms except O and H
3. Balance O by adding H_2O
4. Balance H by adding H^+
5. Balance charge by adding e^-
6. Equalize electrons transferred (multiply)
7. Add half-reactions; cancel common terms

Electrochemical Cells Compared

Feature	Voltaic (Galvanic)	Electrolytic
Energy	Produces electrical energy	Consumes electrical energy
Reaction	Spontaneous ($\Delta G < 0$)	Non-spontaneous
E°_{cell}	Positive (+)	Negative (–)
Anode sign	Negative (–)	Positive (+)
Cathode sign	Positive (+)	Negative (–)
Example	Batteries	Electroplating

Standard Reduction Potentials (E°, 25°C)

Half-Reaction	E° (V)
$F_2 + 2e^- \rightarrow 2F^-$	**+2.87**
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.51
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.23
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
$2H^+ + 2e^- \rightarrow H_2$	0.00
$Fe^{2+} + 2e^- \rightarrow Fe$	–0.44
$Zn^{2+} + 2e^- \rightarrow Zn$	–0.76
$Na^+ + e^- \rightarrow Na$	–2.71

Cell Potential & Gibbs Energy

- **$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$**
- $E^\circ_{cell} > 0$ → spontaneous reaction
- **$\Delta G^\circ = -nFE^\circ_{cell}$** (F = 96,485 C/mol)

Solved Example — Zn-Cu Cell

Cell notation: $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$
E°_{cell} = +0.34 – (–0.76) = **+1.10 V** ✓ (spontaneous)
ΔG° = –2 × 96,485 × 1.10 = **–212.3 kJ**

7.5–7.8 Electrolysis, Faraday's Laws & Applications

Electrolysis — Definition

Using **electrical energy** to drive a **non-spontaneous** redox reaction ($\Delta G > 0$, $E^\circ_{cell} < 0$). Requires an **external power source**. Opposite of a voltaic cell.

Electrolysis of Water

Overall: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$ | Volume ratio $H_2 : O_2$ = **2 : 1**

Electrode	Half-Reaction	Product
Cathode (–)	$4H_2O + 4e^- \rightarrow 2H_2 + 4OH^-$	H_2 gas
Anode (+)	$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$	O_2 gas

Electrolysis of Molten NaCl

Electrode	Half-Reaction	Product
Cathode (–)	$Na^+ + e^- \rightarrow Na$	Na metal
Anode (+)	$2Cl^- \rightarrow Cl_2 + 2e^-$	Cl_2 gas

Electrolysis of NaCl(aq) — Chlor-Alkali Process

Electrode	Half-Reaction	Product
Cathode (–)	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	H_2 gas
Anode (+)	$2Cl^- \rightarrow Cl_2 + 2e^-$	Cl_2 gas

Overall products: H_2, Cl_2, and NaOH solution

Faraday's Laws

- **1st Law:** Mass deposited ∝ charge passed: **$m = (Q \times M) / (n \times F)$**
- **2nd Law:** Same charge deposits masses ∝ equivalent weights (M/n)
- **Variables:** m = mass (g); Q = charge (C) = I × t; M = molar mass (g/mol); n = e^- transferred per ion; F = 96,485 C/mol

Solved Example 1 — Cu Deposition

Cu (M = 63.5 g/mol, n = 2) deposited at **I = 2 A** for **1 hour**:

Step 1: Q = I × t = 2 × 3600 = **7,200 C**

Step 2: m = (Q × M) / (n × F)
m = (7,200 × 63.5) / (2 × 96,485)
m = 457,200 / 192,970 = **2.37 g Cu** ✓

Solved Example 2 — Ag Deposition

Ag (M = 108 g/mol, n = 1) deposited at **I = 3 A** for **2 hours**:

Step 1: Q = I × t = 3 × 7200 = **21,600 C**

Step 2: m = (Q × M) / (n × F)
m = (21,600 × 108) / (1 × 96,485)
m = 2,332,800 / 96,485 = **24.17 g Ag** ✓

Real-World Applications

Application	Details
Electroplating	Coating metals for corrosion protection & aesthetics (jewelry, chrome, silver)
Metal Refining	Purifying Cu: impure anode dissolves; pure Cu deposits at cathode
Chlor-Alkali Industry	Electrolysis of brine → NaOH, Cl_2, H_2; used in plastics, bleach, paper
Batteries	Li-ion (phones/EVs), lead-acid (cars), Ni-MH (hybrids)

🧬 Unit 8: Introduction to Organic Chemistry

The largest unit — carbon's extraordinary bonding versatility gives rise to millions of compounds. Learn the systematic naming, properties, and uses of organic molecules.

Topics Covered

- Why carbon is unique (4 bonds, chains, rings)
- Hydrocarbons: alkanes, alkenes, alkynes, aromatics
- IUPAC naming rules
- Functional groups (alcohols, ethers, aldehydes, ketones, carboxylic acids, esters)
- Fats and oils (triglycerides, saponification)
- Key uses of organic compounds

General Formulas

Alkanes:	C_nH_{2n+2}
Alkenes:	C_nH_{2n}
Alkynes:	C_nH_{2n-2}

8.1–8.3 Hydrocarbons: Classification & Naming

Why Carbon is Special

4 Valence Electrons

Forms exactly **4 covalent bonds** — single, double, or triple — giving unmatched structural flexibility

Chains & Rings

C–C bonds are strong and stable; carbon can chain into **thousands of atoms**, branch, and form rings — basis of all organic life

Hydrocarbon Classification

Class	Formula	Bond Type	Saturation	Examples
Alkane	C_nH_{2n+2}	C–C single	Saturated	Methane CH_4, Ethane C_2H_6, Propane C_3H_8, Butane C_4H_{10}
Alkene	C_nH_{2n}	C=C double	Unsaturated	Ethene C_2H_4, Propene C_3H_6
Alkyne	C_nH_{2n-2}	C≡C triple	Unsaturated	Ethyne C_2H_2, Propyne C_3H_4
Aromatic	C_6H_6 base	Delocalized π ring	Special	Benzene C_6H_6 — resonance structures, delocalized π electrons

Alkane Parent Names (C1–C10)

n	Name	n	Name	n	Name
1	Methane	5	Pentane	8	Octane
2	Ethane	6	Hexane	9	Nonane
3	Propane	7	Heptane	10	Decane
4	Butane				

Common Substituents (Alkyl Groups)

Name	Formula	Structure
Methyl	$-CH_3$	Remove 1 H from methane
Ethyl	$-C_2H_5$	$-CH_2CH_3$
Propyl	$-C_3H_7$	$-CH_2CH_2CH_3$ (straight chain)
Isopropyl	$-C_3H_7$	$-CH(CH_3)_2$ (branched)

IUPAC Naming Rules (Alkanes)

01

Find the longest carbon chain

Parent name: meth-, eth-, prop-, but-, pent-...

02

Number from the nearest branch

Assign **lowest locant numbers** to substituents

03

Name substituents with positions

methyl, ethyl, propyl... prefix with position number

04

Alphabetical order

List multiple substituents alphabetically before the parent name

Example: $CH_3–CH(CH_3)–CH_2–CH_3$ → **2-methylbutane**
Longest chain = 4C (butane); methyl branch at C2

Structural Isomers

Same molecular formula, different structural arrangements → different physical properties.

Formula	Name	Structure
C_4H_{10}	n-Butane	$CH_3–CH_2–CH_2–CH_3$ (straight)
C_4H_{10}	2-Methylpropane (Isobutane)	$CH_3–CH(CH_3)–CH_3$ (branched)

Physical Properties Trend

⬆️ Chain Length

Longer chains → **higher boiling point** (more surface area, stronger London dispersion forces)

⬇️ Branching

More branching → **lower boiling point** (compact shape, less surface contact)

8.4–8.6 Functional Groups, Fats & Uses

Major Functional Groups

Class	Group	Suffix	Example	Common Name	Key Use
Alcohol	–OH	-ol	C_2H_5OH	Ethanol	Antiseptic, solvent, fuel
Ether	–O–	ether	$(C_2H_5)_2O$	Diethyl ether	Anesthetic, solvent
Aldehyde	–CHO	-al	HCHO	Formaldehyde	Preservative, disinfectant
Ketone	C=O (internal)	-one	CH_3COCH_3	Acetone	Nail polish remover, solvent
Carboxylic acid	–COOH	-oic acid	CH_3COOH	Acetic acid	Vinegar, food preservative
Ester	–COO–	-oate	$CH_3COOC_2H_5$	Ethyl ethanoate	Fragrances, flavors, polyesters
Amine	$–NH_2$	-amine	CH_3NH_2	Methylamine	Pharmaceuticals, dyes
Amide	$–CONH_2$	-amide	CH_3CONH_2	Ethanamide	Nylon, proteins

Esterification Reaction

General Equation

Carboxylic acid + Alcohol ⇌ Ester + Water
(reversible; H^+ acid catalyst, heat)

Example

$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$
Ethanoic acid + Ethanol → Ethyl ethanoate

Polymers

Type	Mechanism	Examples
Addition	Alkene monomers open double bond; no byproduct: $nCH_2=CH_2 \rightarrow -(CH_2-CH_2)_n-$	Polyethylene, PVC, polystyrene
Condensation	Bifunctional monomers join; small molecule lost (H_2O or HCl)	Nylon ($–NH_2$ + –COOH), Polyester (–OH + –COOH)

Fats and Oils — Triglycerides

Property	Fats	Oils
Source	Animals	Plants
Bond type	Saturated C–C	Unsaturated C=C
State at room T	Solid	Liquid
Examples	Butter, lard	Olive oil, sunflower oil

Both are **triglycerides** — triesters formed from **glycerol + 3 fatty acid chains** via esterification.

Saponification

Reaction

Fat + NaOH → **Soap** (sodium salt of fatty acid) + **Glycerol**
Base hydrolysis of ester bonds in triglyceride

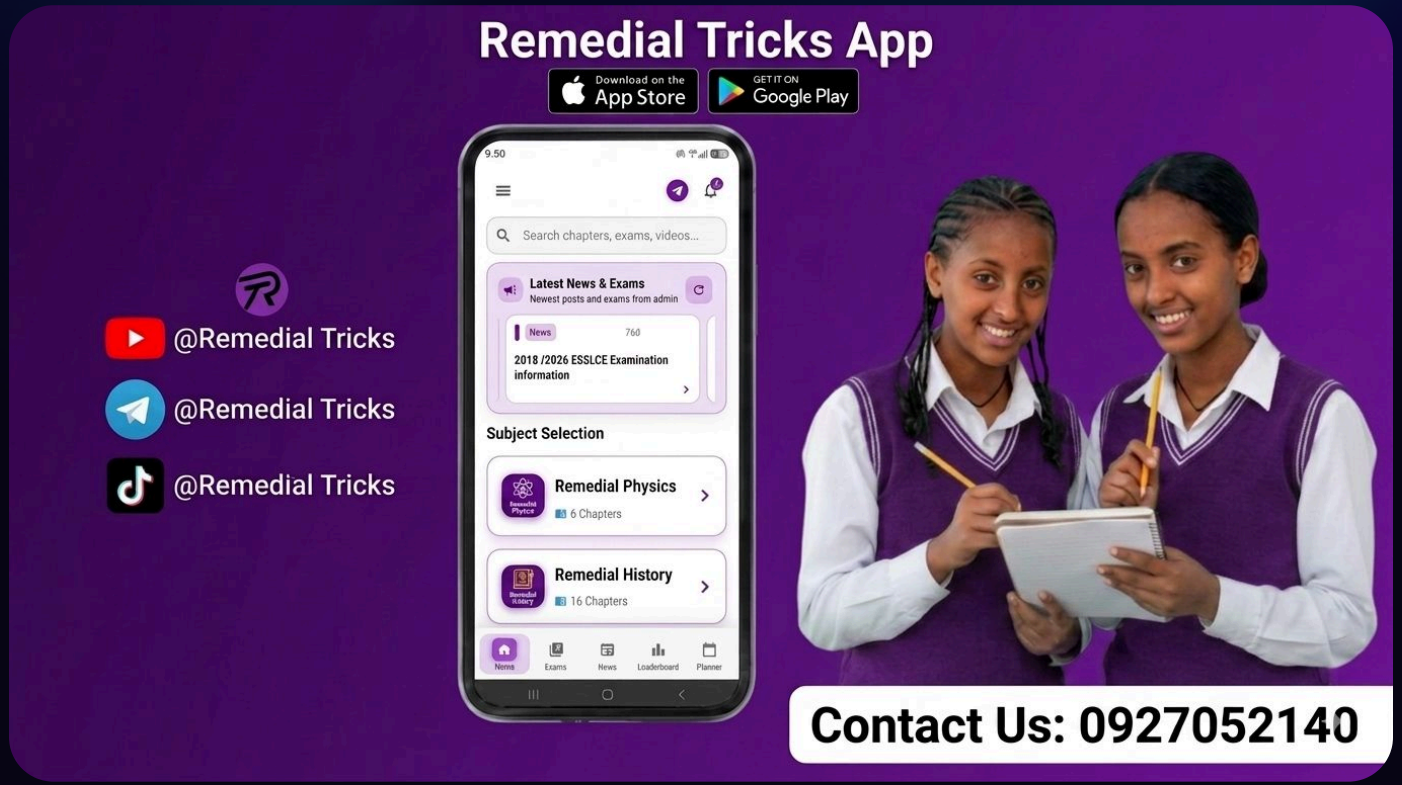
Soap Action — Micelles

Hydrophilic head ($–COO^-Na^+$) faces water
Hydrophobic tail (hydrocarbon chain) faces grease
→ Micelle traps grease; rinsed away by water

Key Uses of Organic Compounds

Category	Examples	Uses
Fuels	Methane, propane, octane	Heating, cooking, transport
Plastics & polymers	Polyethylene, PVC, nylon	Packaging, textiles, construction
Solvents	Ethanol, acetone, ether	Cleaning, pharmaceuticals, labs
Pharmaceuticals	Aspirin (ester), amines	Pain relief, antibiotics, dyes
Food & flavors	Esters, ethanoic acid	Artificial flavors, preservatives, vinegar
Soaps & detergents	Sodium stearate	Cleaning via micelle formation

SECOND ROUND MODEL
(Units 5–8)



REVISION All 8 Units

#	Unit	Key Formula / Concept
1	Atomic Structure	Z, A, quantum numbers (n,l,m_l,m_s), e⁻ config, periodic trends | $E_n = -13.6/n^2$ eV
2	Chemical Bonding	Ionic/covalent/metallic; VSEPR geometry; hybridization (sp, sp^2, sp^3); Bond Order = (bond−antibond)/2
3	States of Matter	PV = nRT; $P_1V_1/T_1 = P_2V_2/T_2$; IMFs: London → dipole → H-bond; solid types; phase diagrams
4	Solutions	M = mol/L; $M_1V_1 = M_2V_2$; solubility rules; "like dissolves like"; Henry's Law C = kP
5	Stoichiometry	Balance → moles → mole ratio → mass; ΔH = ΣΔH_f(prod) − ΣΔH_f(react)
6	Equilibrium	Kc = [P]^p/[R]^r; Le Chatelier; Haber: $N_2+3H_2 \rightleftharpoons 2NH_3$, 400–500°C, 150–300 atm, Fe
7	Electrochemistry	OIL RIG; E°cell = E°cat − E°an; m = QM/nF; Q = It; F = 96485 C/mol
8	Organic Chemistry	$C_nH_{2n+2}/C_nH_{2n}/C_nH_{2n-2}$; IUPAC naming; –OH, –CHO, –COOH, –COO–; saponification

Final Encouragement

> 🌟 You have covered **70 hours** of chemistry. Review your notes, practice past papers, and believe in yourself. You are more prepared than you think!

Last-Minute Strategy

01

Review all 8 formula sheets

One unit per study session

02

Practice all worked examples

Redo them without looking at the answer

03

Attempt the Final Model Exam

Simulate exam conditions: 3 hours, no notes

🏆 THIRD ROUND MODEL

UNIT(1-8)

