CHEMISTRY · UNITS 1–9 · QUICK REFERENCE

General Chemistry

Grades 9–12+ · 8 pages

Measurement, atoms and electrons, trends, bonding and shape, forces, naming, the mole and stoichiometry, reactions, solutions, gases, energy, rates, equilibrium, acids and bases, redox and nuclear chemistry.

1 Measurement · 2 Atoms · 3 Electrons · 4 Trends · 5 Bonding · 6 Shape & polarity · 7 IMFs & phases · 8 Naming · 9 Mole & stoichiometry · 10 Reactions · 11 Solutions · 12 Gases · 13 Thermochemistry · 14 Rates & equilibrium · 15 Acids & bases · 16 Redox & nuclear

1 MEASUREMENT, SIG FIGS & UNITS

SI prefixes

Prefix	Symbol	Factor
giga	G	10^{9}
mega	M	10^{6}
kilo	k	10^{3}
centi	c	10^{-2}
milli	m	10^{-3}
micro	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}

SI base kg (mass), m (length), s (time), K (temperature), mol (amount)

Accuracy how close to the true value

Precision how close repeated results are to each other

Rounding answers

- × and ÷: keep the **fewest sig figs** of any factor.
- + and −: keep the **fewest decimal places**.
- Round once, at the end.

EX 4.56 × 1.4 = 6.384 → **6.4**; 12.11 + 3.2 = 15.31 → **15.3**

Counting significant figures

- Nonzero digits always count.
- Zeros **between** nonzero digits count: 3**0**4 has 3.
- Leading zeros never count: 0.00**45** has 2.
- Trailing zeros count only with a decimal point: 45.0**0** has 4; 4500 is unclear (write 4.5×10^{3} or 4.500×10^{3}).
- Exact numbers (counted items, defined units) have unlimited sig figs.

MEMORY TRICK

Middle yes, front no, end needs a dot

Zeros in the **middle** always count, zeros in **front** never do, zeros at the **end** count only when the number shows a decimal **dot**.

Count them WORKED

0.004 50 → **3**; 30.040 → **5**; 1.200×10^{3} → **4**

Percent error WORKED

measured 2.62 g/cm³, accepted 2.70 g/cm³

$\frac{|2.62 - 2.70|}{2.70} \times 100 = \mathbf{3.0\ \%}$

Dimensional analysis

Multiply by conversion factors (each equal to 1) until only the wanted unit is left.

$$\text{given} \times \frac{\text{wanted unit}}{\text{given unit}}$$

Convert 65 mi/h to m/s WORKED

$65 \frac{\text{mi}}{\text{h}} \times \frac{1609\ \text{m}}{1\ \text{mi}} \times \frac{1\ \text{h}}{3600\ \text{s}} = \mathbf{29\ m/s}$

1 mi = 1609.344 m

Density WORKED

12.0 g stone raises water from 20.0 mL to 24.4 mL

$d = \frac{m}{V} = \frac{12.0\ \text{g}}{4.4\ \text{mL}} = \mathbf{2.7\ g/mL}$ *4.4 has 2 sig figs*

Temperature

$$K = {}^{\circ}C + 273.15$$

$$^{\circ}F = 1.8({}^{\circ}C) + 32$$

EX 37.0 °C = **310.2 K** = **98.6 °F**

2 ATOMS, ISOTOPES & IONS

Subatomic particles

Particle	Charge	Mass (u)	Where
proton	+1	1.007	nucleus
neutron	0	1.009	nucleus
electron	−1	0.000 549	outside

MEMORY TRICK

P picks, N nudges, E earns

Protons **pick** the element, **n**eutrons **nudge** the mass, **e**lectrons **earn** the charge.

Nuclide symbol

Z = atomic number (protons); A = mass number (protons + neutrons); N = neutrons.

$$^{A}_{Z}X: A = Z + N$$

Count the particles WORKED

$^{56}Fe^{3+}$: Z = 26

26 p, 30 n, 23 e⁻ *lost 3 e⁻*

Isotopes same Z, different N (^{12}C, ^{13}C, ^{14}C)

Ion unequal p and e⁻; cation (+) lost e⁻, anion (−) gained e⁻

Atomic mass unit 1 u = $\frac{1}{12}$ the mass of a ^{12}C atom

Average atomic mass

The weighted average printed on the periodic table.

$$\Sigma\ (\text{fraction} \times \text{isotope mass})$$

Chlorine WORKED

75.78 % ^{35}Cl (34.969 u), 24.22 % ^{37}Cl (36.966 u)

0.7578(34.969) + 0.2422(36.966) = **35.45 u**

Find the abundance WORKED

Cu = 63.546 u from ^{63}Cu (62.930 u) and ^{65}Cu (64.928 u)

62.930x + 64.928(1 − x) = 63.546 → x = **0.6917**

69.17 % ^{63}Cu

Models of the atom

Year	Scientist	Model
1808	Dalton	solid indivisible spheres
1897	Thomson	electrons in positive matter
1911	Rutherford	tiny dense nucleus
1913	Bohr	electrons in energy levels
1926	Schrödinger	orbitals (probability)

3 ELECTRONS, LIGHT & QUANTUM NUMBERS

Light

$c = 2.998 \times 10^8$ m/s; $h = 6.626 \times 10^{-34}$ J·s.

$$c = \lambda\nu$$

$$E = h\nu = \frac{hc}{\lambda}$$

Hydrogen's red line, 656 nm WORKED

$\nu = \frac{2.998 \times 10^8}{656 \times 10^{-9}} = \mathbf{4.57 \times 10^{14}\ Hz}$

$E = h\nu = \mathbf{3.03 \times 10^{-19}\ J}$ *per photon*

Bohr energy levels (hydrogen)

An electron dropping from a high level to a low one gives off a photon of exactly ΔE.

$$E_n = -\frac{2.18 \times 10^{-18}\ J}{n^2}$$

EX $n = 3 \to 2$: $\Delta E = 2.18 \times 10^{-18}(\frac{1}{4} - \frac{1}{9}) = \mathbf{3.03 \times 10^{-19}\ J}$, the same 656 nm red line

Electromagnetic spectrum (low → high energy)

radio < microwave < infrared < visible (red → violet) < ultraviolet < X-ray < gamma

Longer wavelength means lower frequency and lower energy.

Quantum numbers

Number	Values	Tells
n	1, 2, 3, …	energy level, size
l	0 to n − 1 (s, p, d, f)	sublevel shape
m_l	−l to +l	orbital orientation
m_s	+½ or −½	spin

Which set is impossible? WORKED

(2, 1, 0, +½) ✓; (3, 2, −2, −½) ✓

(2, 2, 0, +½) **✗** *l must be less than n*

Sublevels s 1 orbital (2 e⁻), p 3 (6), d 5 (10), f 7 (14)

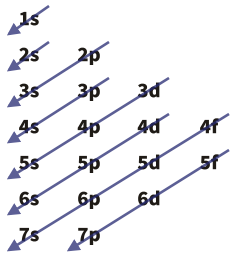
Level n holds $2n^2$ electrons: 2, 8, 18, 32

MEMORY TRICK **2, 6, 10, 14: add four**

Each new sublevel holds **four more** electrons: s 2, p 6, d 10, f 14.

Filling rules

- **Aufbau:** fill the lowest energy first (diagonal chart).
- **Pauli:** an orbital holds 2 e⁻ of opposite spin; no two e⁻ share all four numbers.
- **Hund:** in equal orbitals, spread out one per orbital (same spin) before pairing.



Follow each arrow from its top end: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p.

Sulfur (16 e⁻) WORKED

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ = **[Ne] $3s^2\ 3p^4$**

3p: ↑↓ ↑ ↑ → **2 unpaired** *Hund*

Bromine (35 e⁻) WORKED

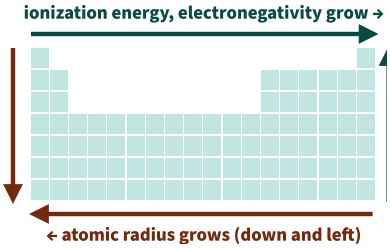
[Ar] $4s^2\ 3d^{10}\ 4p^5$

outer level n = 4: 2 + 5 = **7 valence e⁻** *group 17*

⚠ Watch for

- Cr = [Ar] $4s^1\ 3d^5$ and Cu = [Ar] $4s^1\ 3d^{10}$ (half- and full-d stability).
- Transition-metal ions lose **4s before 3d**: Fe^{2+} = [Ar] $3d^6$.
- Full table with every configuration block: Periodic Table sheet.

4 PERIODIC TRENDS



Pull on electrons rises to the upper right; size grows to the lower left.

Trend table

Property	Across →	Down ↓
atomic radius	smaller	larger
ionization energy	higher	lower
electronegativity	higher	lower
electron affinity	more negative (mostly)	less negative
metallic character	less	more

MEMORY TRICK **Up-right holds tight**

Atoms toward the **upper right** hold electrons **tightest**: smallest radius, highest ionization energy and electronegativity (noble gases aside).

Families

Group	Family	Valence e⁻	Ion
1	alkali metals	1	1+
2	alkaline earth metals	2	2+
3–12	transition metals	—	varies
15	nitrogen family	5	3−
16	chalcogens	6	2−
17	halogens	7	1−
18	noble gases	8 (He 2)	none

Effective nuclear charge

$$Z_{eff} \approx Z - \text{core electrons}$$

EX Na: 11 − 10 = **1**; Cl: 17 − 10 = **7**: Cl holds its electrons far tighter.

Rank the radius WORKED

K, Na, Mg, Cl

K > Na > Mg > Cl *down is bigger; right is smaller*

Rank electronegativity WORKED

C, N, O, F

F (3.98) > O (3.44) > N (3.04) > C (2.55)

Ionization energy energy to remove an e⁻ from a gas atom

Electronegativity pull on shared electrons in a bond (Pauling scale; F = 3.98)

Electron affinity energy change when a gas atom gains an e⁻

Ion size

- Cations are **smaller** than their atoms (lost a shell or e⁻).
- Anions are **larger** (extra e⁻ repel).
- Same electron count: more protons → smaller: $O^{2-} > F^- > Na^+ > Mg^{2+}$.

Read successive ionization energies WORKED

Mg: 738, 1451, 7733 kJ/mol

huge jump after the 2nd → **2 valence e⁻**

group 2

⚠ Exceptions

- Ionization energy dips at B (below Be) and O (below N).
- Noble gases have no Pauling electronegativity (most form no bonds).

5 BONDING & LEWIS STRUCTURES

Bond type by ΔEN

ΔEN	Bond	Example
< 0.4	nonpolar covalent	C–H 0.35
0.4 – 1.7	polar covalent	H–Cl 0.96
> 1.7	mostly ionic	Na–Cl 2.23

Three kinds

Bond	Made of	Properties
ionic	metal + nonmetal ions	high melting, brittle; conducts melted or dissolved
covalent	nonmetals sharing e^-	molecules; lower melting
metallic	metal cations in an e^- sea	conducts, malleable, shiny

Drawing a Lewis structure

- **Count** valence e^- (add for − charge, subtract for +).
- **Connect** atoms with single bonds; least electronegative in the center, H on the outside.
- **Complete** octets on outer atoms.
- Put leftovers on the **center** atom.
- **Check:** center short of 8? Turn lone pairs into double or triple bonds.

MEMORY TRICK **Five Cs**
Count, **C**onnect, **C**omplete, **C**enter, **C**heck.

CO_2 WORKED

4 + 2(6) = **16 e^-**
O=C=O, two lone pairs on each O *all octets*

NH_3 WORKED

5 + 3(1) = **8 e^-**
3 N–H bonds + **1 lone pair** on N

Formal charge

V valence e^-, L lone-pair e^-, B bonding e^-. Best structure: charges near zero, any − on the more electronegative atom.

$$FC = V - L - \frac{B}{2}$$

Carbonate, CO_3^{2-} WORKED

4 + 3(6) + 2 = **24 e^-**; one C=O, two C–O
single-bonded O: 6 − 6 − 1 = **−1**; C and double-bonded O: **0**
3 resonance forms

Octet exceptions

- H holds 2; Be often 4; B often 6 (BF_3).
- Period 3 and beyond can exceed 8: PCl_5 (10), SF_6 (12).
- Odd electron counts: NO (11 e^-).

Bond energies estimate ΔH WORKED

$H_2 + Cl_2 \rightarrow 2\ HCl$; H–H 436, Cl–Cl 243, H–Cl 431 kJ/mol
ΔH ≈ broken − formed = 436 + 243 − 2(431) = **−183 kJ**
exothermic

Strength of ionic bonds (lattice energy)

- Bigger ion charges → stronger attraction.
- Smaller ions → closer together → stronger.

So MgO (2+, 2−) melts far higher than NaCl (1+, 1−).

Bond order single 1, double 2, triple 3
Higher order shorter and stronger: C≡C < C=C < C–C in length

Resonance two or more valid Lewis structures; the real molecule is their average (O_3, CO_3^{2-})
Coordinate bond both shared electrons come from one atom (NH_4^+)

6 SHAPE, HYBRIDS & POLARITY

VSEPR shapes (domains = bonded atoms + lone pairs on the center atom; a double or triple bond counts once)

Domains	Lone	Shape	Angle	Hybrid	Example
2	0	linear	180°	sp	CO_2
3	0	trigonal planar	120°	sp^2	BF_3
3	1	bent	< 120°	sp^2	SO_2
4	0	tetrahedral	109.5°	sp^3	CH_4
4	1	trigonal pyramidal	≈ 107°	sp^3	NH_3
4	2	bent	≈ 104.5°	sp^3	H_2O
5	0	trigonal bipyramidal	90°, 120°	sp^3d	PCl_5
5	1	seesaw	< 90°, < 120°	sp^3d	SF_4
5	2	T-shaped	< 90°	sp^3d	ClF_3
5	3	linear	180°	sp^3d	XeF_2
6	0	octahedral	90°	sp^3d^2	SF_6
6	1	square pyramidal	< 90°	sp^3d^2	BrF_5
6	2	square planar	90°	sp^3d^2	XeF_4

Is the molecule polar?

- No polar bonds → nonpolar.
- Polar bonds in a **symmetric** shape cancel → nonpolar (CO_2, BF_3, CCl_4).
- Polar bonds in a **lopsided** shape (lone pairs on the center or mixed outer atoms) → polar (H_2O, NH_3, $CHCl_3$).

MEMORY TRICK **Lone pairs push harder**
A lone pair spreads wider than a bond, squeezing the angles: 109.5° (CH_4) → 107° (NH_3) → 104.5° (H_2O).

Polar or not? WORKED

CCl_4: tetrahedral, four equal pulls → **nonpolar**
H_2O: bent → **polar**
SO_2: bent (1 lone pair) → **polar**

σ bond head-on overlap; every single bond
π bond side-by-side overlap; a double bond is 1σ + 1π, a triple 1σ + 2π

Count σ and π in C_2H_4 WORKED

4 C–H + 1 C–C σ = **5 σ**
C=C adds **1 π** *each C is sp^2*

Predict XeF_4 WORKED

Xe: 8 valence e^-; 4 bonds use 4 → 2 lone pairs
6 domains, 2 lone → **square planar**, **nonpolar**
lone pairs opposite

7 INTERMOLECULAR FORCES & PHASES

Intermolecular forces

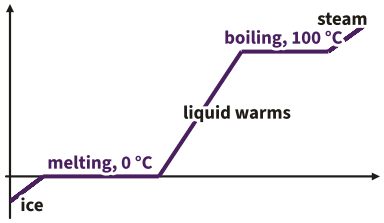
Force	Between	Strength
London dispersion	all molecules; grows with size (more e^-)	weakest per contact
dipole–dipole	polar molecules	moderate
hydrogen bond	H on N, O or F → lone pair on N, O or F	strongest IMF
ion–dipole	ions and polar molecules (salt in water)	stronger still

MEMORY TRICK **H bonds need F, O, N**
Hydrogen bonds form only when H sits on **F, O or N**. (Traditional "H-FON".)

Stronger IMFs mean
- higher boiling and melting points
- lower vapor pressure (less evaporation)
- higher surface tension and viscosity

Rank boiling points WORKED
CH_4 (−162 °C) < HCl (−85 °C) < H_2O (100 °C)
dispersion < dipole < **hydrogen bonding**

Halogens at room temperature WORKED
F_2, Cl_2 gases; Br_2 liquid; I_2 solid
more electrons → stronger dispersion



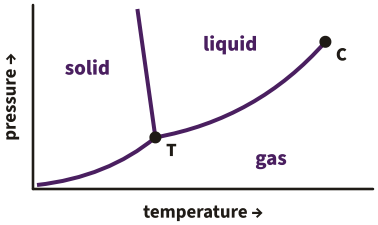
Heating curve for water: temperature holds steady while the phase changes.

Phase changes

Change	Name	Heat
solid → liquid	melting (fusion)	absorbed
liquid → gas	vaporization	absorbed
solid → gas	sublimation	absorbed
gas → liquid	condensation	released
liquid → solid	freezing	released
gas → solid	deposition	released

Melt and warm 10.0 g of ice at 0 °C to 25.0 °C WORKED
$q = m\Delta H_{fus} + mc\Delta T$ *ΔH_{fus} = 334 J/g*
$= 10.0(334) + 10.0(4.184)(25.0) =$ **4.39 kJ**
3340 + 1046 J

Boiling point vapor pressure equals the outside pressure
Triple point all three phases together
Critical point above it, liquid and gas merge



Phase diagram (water-like, not to scale): lines are where two phases coexist; T triple point, C critical point.

Boiling and pressure
- Lower outside pressure (high altitude) → water boils below 100 °C.
- A pressure cooker raises the pressure, so water boils hotter and food cooks faster.

8 NAMING COMPOUNDS

Ionic compounds
- Cation name, then anion name (-ide for one-element anions).
- Metals with more than one charge get a Roman numeral (Stock system).
- Polyatomic ions keep their names; list: Periodic Table sheet.

Name it WORKED
$CaCl_2$ → **calcium chloride**
Fe_2O_3 → **iron(III) oxide**
2 Fe = +6 balances 3 O = −6
$(NH_4)_2SO_4$ → **ammonium sulfate**
$Cu(NO_3)_2$ → **copper(II) nitrate**

Writing the formula
- Total charge must be zero.
- Cross the charge numbers to subscripts, then reduce.
- Use parentheses for 2 or more polyatomic ions.

EX $Al^{3+} + O^{2-} \rightarrow$ **Al_2O_3**; $Ca^{2+} + PO_4^{3-} \rightarrow$ **$Ca_3(PO_4)_2$**

MEMORY TRICK **Zero is the goal**
Every ionic formula adds to **zero** charge: check it by multiplying each charge by its subscript.

Covalent prefixes (two nonmetals)
1 mono, 2 di, 3 tri, 4 tetra, 5 penta, 6 hexa, 7 hepta, 8 octa, 9 nona, 10 deca

Most-used polyatomic ions (full list: Periodic Table sheet)

Ion	Name	Ion	Name
NH_4^+	ammonium	OH^-	hydroxide
NO_3^-	nitrate	SO_4^{2-}	sulfate
CO_3^{2-}	carbonate	PO_4^{3-}	phosphate
HCO_3^-	hydrogen carbonate	$C_2H_3O_2^-$	acetate

Write the formula WORKED
iron(III) sulfate → **$Fe_2(SO_4)_3$** *2(+3) + 3(−2) = 0*
sodium hydrogen carbonate → **$NaHCO_3$**
baking soda
dinitrogen pentoxide → **N_2O_5**

Covalent names WORKED
N_2O_4 → **dinitrogen tetroxide** *drop "a" before o*
CO → **carbon monoxide**
no mono- on the first element
PCl_5 → **phosphorus pentachloride**

Acids (in water)

Anion ends in	Acid name	Example
-ide	hydro___ic acid	HCl hydrochloric acid
-ate	___ic acid	HNO_3 nitric, H_2SO_4 sulfuric
-ite	___ous acid	HNO_2 nitrous, H_2SO_3 sulfurous

MEMORY TRICK **I ate something icky**
-ate → **-ic**; **-ite** → **-ous** ("the sprITE was delicOUS"). (Traditional.)

Hydrate water in the crystal: $CuSO_4 \cdot 5H_2O$, copper(II) sulfate pentahydrate
Diatomic H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2 as elements
Common names H_2O water, NH_3 ammonia, CH_4 methane, H_2O_2 hydrogen peroxide

9 THE MOLE & STOICHIOMETRY

Mole	6.022×10^{23} particles (Avogadro's number)
Molar mass	grams per mole; add the atomic weights
Molar volume	22.4 L per mole of gas at STP (0 °C, 1 atm)

The mole map

- grams ÷ molar mass → **moles** × molar mass → grams
- particles ÷ 6.022×10^{23} ↔ **moles** × 6.022×10^{23}
- liters of gas (STP) ÷ 22.4 ↔ **moles**
- molarity × liters of solution → **moles**

MEMORY TRICK

Every road runs through moles

Never jump straight from grams of one thing to grams of another: go **through moles**.

Molar mass of $Ca(NO_3)_2$ WORKED

40.078 + 2(14.007) + 6(15.999)

= **164.09 g/mol**

25.0 g of water WORKED

25.0 ÷ 18.015 = **1.39 mol**

1.39 mol × 6.022×10^{23} = **8.36×10^{23} molecules**

Percent composition

Use the element's total mass in one mole of the compound.

$$\% = \frac{\text{element's mass}}{\text{molar mass}} \times 100$$

EX O in H_2O: $\frac{15.999}{18.015} \times 100$ = **88.81 %**

Empirical formula from percents

- Treat percent as grams (100 g sample).
- Grams → moles of each element.
- Divide by the smallest; round or multiply to whole numbers.
- Molecular formula: multiply by $\frac{\text{molar mass}}{\text{empirical mass}}$.

40.0 % C, 6.71 % H, 53.3 % O; 180.16 g/mol WORKED

C 3.33, H 6.66, O 3.33 mol → 1 : 2 : 1 → **CH_2O**

180.16 ÷ 30.026 = 6 → **$C_6H_{12}O_6$**

Steps

- Balance the equation.
- Given → **moles**.
- Multiply by the **mole ratio** from coefficients.
- Moles → asked-for unit.

Water from 10.0 g H_2 WORKED

$2\ H_2 + O_2 \rightarrow 2\ H_2O$

10.0 ÷ 2.016 = 4.96 mol $H_2 \times \frac{2\ H_2O}{2\ H_2}$ = 4.96 mol H_2O

× 18.015 = **89.4 g H_2O**

Limiting reactant

- Work out the product from **each** reactant.
- The one giving **less** product is limiting; that amount is the theoretical yield.
- The other is in excess: subtract what reacted to find what is left.

MEMORY TRICK **The pantry rule**

The ingredient that **runs out first** sets the size of the batch, no matter how much of the rest you have.

Ammonia: 28.0 g N_2 + 9.0 g H_2 WORKED

$N_2 + 3\ H_2 \rightarrow 2\ NH_3$

N_2: 0.9995 mol → 1.999 mol NH_3; H_2: 4.46 mol → 2.98 mol NH_3

N_2 limits: **34.0 g NH_3** *× 17.031 g/mol*

H_2 left: 4.46 − 3.00 = 1.47 mol = **3.0 g**

Percent yield

$$\%\ \text{yield} = \frac{\text{actual}}{\text{theoretical}} \times 100$$

EX got 25.5 g NH_3 of 34.0 g possible: **75.0 %**

10 REACTIONS & NET IONIC EQUATIONS

Balancing

- Change **coefficients** only, never subscripts.
- Balance metals, then nonmetals, then H, then O last.
- Treat an unchanged polyatomic ion as one unit.

EX $C_3H_8 + \mathbf{5}\ O_2 \rightarrow \mathbf{3}\ CO_2 + \mathbf{4}\ H_2O$; $\mathbf{4}\ Al + \mathbf{3}\ O_2 \rightarrow \mathbf{2}\ Al_2O_3$

Five reaction types

Type	Pattern	Example
synthesis	A + B → AB	$2\ Na + Cl_2 \rightarrow 2\ NaCl$
decomposition	AB → A + B	$2\ H_2O_2 \rightarrow 2\ H_2O + O_2$
single replacement	A + BC → AC + B	$Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$
double replacement	AB + CD → AD + CB	$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
combustion	fuel + O_2 → CO_2 + H_2O	$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

Solubility in water

Ions	Soluble?	Exceptions
Group 1, NH_4^+, NO_3^-, $C_2H_3O_2^-$	always	none
Cl^-, Br^-, I^-	yes	Ag^+, Pb^{2+}, Hg_2^{2+}
SO_4^{2-}	yes	Ba^{2+}, Pb^{2+}, Ca^{2+}, Sr^{2+}
CO_3^{2-}, PO_4^{3-}, S^{2-}	no	with Group 1 or NH_4^+
OH^-	no	Group 1, NH_4^+; Ba^{2+}, Sr^{2+}, Ca^{2+} partly

Net ionic equation

- Write the balanced molecular equation with states.
- Split strong electrolytes (aq) into ions; keep solids, liquids, gases and weak acids whole.
- Cross out **spectator ions** (same on both sides).

MEMORY TRICK **Spectators just watch**

Ions that look the same on both sides **watch the game** and leave the final equation.

Yellow precipitate WORKED

$Pb(NO_3)_2(aq) + 2\ KI(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$

net: **$Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$**

K^+, NO_3^- watch

Any strong acid + strong base WORKED

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Activity series (most → least active)

Li > K > Ba > Ca > Na > Mg > Al > Zn > Fe > Ni > Sn > Pb > (H) > Cu > Ag > Au

A metal replaces any ion **below** it: $Cu + 2\ Ag^+ \rightarrow Cu^{2+} + 2\ Ag$ works; $Cu + Zn^{2+}$ does not.

11 SOLUTIONS & CONCENTRATION

Solute / solvent what dissolves / what does the dissolving

Saturated holds all it can at that temperature

Like dissolves like polar with polar (salt, sugar in water), nonpolar with nonpolar (oil in hexane)

Electrolyte makes ions in water: strong (NaCl, HCl), weak (acetic acid), non (sugar)

Molarity

$$M = \frac{\text{mol solute}}{\text{L solution}}$$

Dilution

$$M_1V_1 = M_2V_2$$

Mass percent WORKED

15.0 g sugar in 135.0 g water

$\frac{15.0}{150.0} \times 100 =$ **10.0 %** *solution mass = 150.0 g*

5.85 g NaCl in 250.0 mL WORKED

5.85 ÷ 58.44 = 0.100 mol

0.100 mol ÷ 0.2500 L = **0.400 M**

MEMORY TRICK **MV makes moles**

Molarity × volume (L) = **moles**: the bridge for every solution problem.

500. mL of 1.50 M from 12.0 M HCl WORKED

$V_1 = \frac{1.50 \times 500.}{12.0} =$ **62.5 mL** *add acid to water*

Titration WORKED

18.40 mL of 0.1000 M NaOH neutralizes 25.00 mL HCl

mol NaOH = 1.840×10^{-3} = mol HCl *1 : 1*

$[HCl] = \frac{1.840 \times 10^{-3}}{0.02500\ L} =$ **0.07360 M**

Molality m = mol solute ÷ kg solvent

Mass percent (g solute ÷ g solution) × 100

ppm (g solute ÷ g solution) × 10^6

Colligative properties

Water: $K_f = 1.86$, $K_b = 0.512$ °C·kg/mol; i = particles per formula unit.

$$\Delta T_f = i\,K_f\,m$$

$$\Delta T_b = i\,K_b\,m$$

EX 1.00 m NaCl (i = 2) freezes at **−3.72 °C**

Molality and freezing WORKED

18.0 g glucose (180.16 g/mol) in 500. g water

0.0999 mol ÷ 0.500 kg = **0.200 m** → ΔT_f = 1.86(0.200) = **0.372 °C**

i = 1

Ions in solution WORKED

0.10 M $CaCl_2$ → $[Ca^{2+}]$ = **0.10 M**, $[Cl^-]$ = **0.20 M**

2 Cl^- per formula

Solubility and temperature

- Most solids dissolve better when hot.
- Gases dissolve better when **cold** and under higher pressure (an opened soda goes flat).

12 GAS LAWS

Before you start

- Temperature in **kelvin**, always.
- 1 atm = 760 mmHg = 760 torr = 101.325 kPa.
- STP: 0 °C (273.15 K) and 1 atm.

The gas laws

Law	Equation	Held constant
Boyle	$P_1V_1 = P_2V_2$	n, T
Charles	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	n, P
Gay-Lussac	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	n, V
Avogadro	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$	P, T
Combined	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	n

MEMORY TRICK

Boyle is the one that flips

Squeeze a gas and pressure **rises** as volume **falls**: the only inverse pair. Every other pair moves together.

Boyle WORKED

2.0 L at 1.0 atm → 0.50 atm

$V_2 = \frac{1.0 \times 2.0}{0.50} =$ **4.0 L**

Charles WORKED

3.00 L at 27 °C → 127 °C

$V_2 = 3.00 \times \frac{400.15}{300.15} =$ **4.00 L** *kelvin!*

Combined: to STP WORKED

2.50 L at 25 °C, 1.00 atm → 0 °C, 1.00 atm

$V_2 = 2.50 \times \frac{273.15}{298.15} =$ **2.29 L**

Ideal gas law

R = 0.08206 L·atm/(mol·K) = 8.314 J/(mol·K).

$$PV = nRT$$

Find P WORKED

0.500 mol in 10.0 L at 25 °C

$P = \frac{0.500 \times 0.08206 \times 298.15}{10.0} =$ **1.22 atm**

Molar mass from density WORKED

$M = \frac{dRT}{P}$: 1.25 g/L at STP

1.25 × 0.08206 × 273.15 ÷ 1 = **28.0 g/mol**

N_2 or CO

Dalton: partial pressures

$$P_{total} = P_1 + P_2 + \ldots$$

EX gas over water: 750. − 23.8 (water vapor at 25 °C) = **726 mmHg** dry gas

Graham: effusion

$$\frac{rate_1}{rate_2} = \sqrt{\left(\frac{M_2}{M_1}\right)}$$

EX He vs O_2: √(32.00 ÷ 4.003) = **2.83** times faster

Gas stoichiometry WORKED

Zn + 2 HCl → $ZnCl_2$ + H_2; 5.00 g Zn at 25 °C, 1.00 atm

5.00 ÷ 65.38 = 0.0765 mol H_2

$V = \frac{nRT}{P} = 0.0765(0.08206)(298.15) =$ **1.87 L**

Kinetic-molecular theory

- Particles are tiny compared with the space between them.
- They move randomly and collide elastically.
- Average kinetic energy ∝ kelvin temperature.
- Real gases stray from ideal at **high P** (molecules take up room) and **low T** (attractions matter).

13 THERMOCHEMISTRY

MEMORY TRICK **Exo exits**
In an **exo**thermic change heat **exits** the system, so ΔH is negative and the surroundings warm up.

Heat
Exothermic: releases heat, ΔH < 0. Endothermic: absorbs heat, ΔH > 0. Water: c = 4.184 J/(g·°C).

$$q = mc\Delta T$$

EX 150. g water, 22.0 → 35.5 °C: q = 150.(4.184)(13.5) = **8.47 kJ**

Coffee-cup calorimeter WORKED
50.0 g metal at 100.0 °C into 100.0 g water at 20.0 °C; final 23.7 °C

heat lost by metal = heat gained by water

$$c = \frac{100.0(4.184)(3.7)}{50.0(76.3)} = \mathbf{0.41\ J/(g\cdot°C)}$$

Hess's law: C + ½O₂ → CO WORKED
$C + O_2 \rightarrow CO_2$ ΔH = −393.5 kJ

$CO_2 \rightarrow CO + \frac{1}{2}O_2$ ΔH = +283.0 kJ *reversed*

add the steps and their ΔH: **ΔH = −110.5 kJ**

reverse a step → flip the sign

From enthalpies of formation
$\Delta H°_f$ of an element in its standard state = 0.

$$\Delta H°_{rxn} = \Sigma n\Delta H°_f(\text{products}) - \Sigma n\Delta H°_f(\text{reactants})$$

Burning methane WORKED
$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O(l)$

ΔH°f: −74.8, 0, −393.5, −285.8

−393.5 + 2(−285.8) − (−74.8) = **−890.3 kJ**

16.0 g CH_4 (0.997 mol) releases **888 kJ**

Gibbs free energy
ΔG < 0: spontaneous (product-favored). ΔS > 0: more disorder (gas made, solid dissolves).

$$\Delta G = \Delta H - T\Delta S$$

Sign chart

ΔH	ΔS	Spontaneous?
−	+	always
+	−	never
−	−	at low T
+	+	at high T

Ammonia at 298 K WORKED
ΔH = −92.2 kJ, ΔS = −198.7 J/K

ΔG = −92.2 − 298(−0.1987) = **−33.0 kJ**

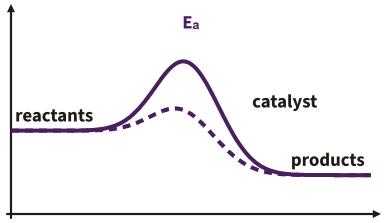
spontaneous

switches at $T = \frac{\Delta H}{\Delta S} \approx$ **464 K**

14 REACTION RATES & EQUILIBRIUM

Collision theory
- Particles must **collide**,
- with the right **orientation**,
- and with energy ≥ the **activation energy** E_a.

MEMORY TRICK **SCAT speeds it up**
Surface area, **C**oncentration, **A**ctivation helpers (catalysts), **T**emperature: raise any of them and the rate rises.



Exothermic reaction: products sit lower (ΔH < 0). A catalyst (dashed) lowers E_a but not ΔH.

Rate law
Orders m and n come from experiments, not from coefficients. Overall order = m + n.

$$\text{rate} = k[A]^m[B]^n$$

Orders from data WORKED
double [A], rate × 4 → **2nd order in A**

double [B], rate unchanged → **0 order in B**

rate = $k[A]^2$: 0.010 M/s at [A] = 0.10 M → **k = 1.0 $M^{-1}s^{-1}$**

By order

Order	Rate law	Units of k	Half-life
0	k	M/s	$[A]_0 \div 2k$
1	k[A]	s^{-1}	0.693 ÷ k
2	$k[A]^2$	$M^{-1}s^{-1}$	$1 \div (k[A]_0)$

First-order half-life WORKED
$k = 0.0231\ min^{-1} \rightarrow t_{½} = \frac{0.693}{0.0231} =$ **30.0 min**

after 90.0 min (3 half-lives): **⅛** remains

Mechanism the series of elementary steps

Rate-determining step the slowest step sets the rate

Catalyst used, then regenerated; offers a lower-E_a path

Equilibrium constant
For aA + bB ⇌ cC + dD. Leave out pure solids and liquids. K ≫ 1 favors products; K ≪ 1 favors reactants.

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

Q versus K
- Q uses the current amounts in the K expression.
- Q < K: shifts **right** (makes products).
- Q > K: shifts **left**.
- Q = K: at equilibrium.

ICE table: $H_2 + I_2 \rightleftharpoons 2\ HI$, K = 49.0 WORKED
start 1.00 M H_2, 1.00 M I_2; change −x, −x, +2x

$\frac{(2x)^2}{(1.00-x)^2} = 49.0 \rightarrow \frac{2x}{1.00-x} = 7.00 \rightarrow x = 0.778$

[HI] = 1.56 M; [H_2] = [I_2] = 0.222 M

Le Chatelier's principle

Stress	Shift
add reactant or remove product	→ right
add product or remove reactant	← left
raise pressure (shrink volume)	toward fewer gas moles
raise temperature	toward the endothermic side (K changes)
add a catalyst	no shift; just faster

MEMORY TRICK
Push it and it pushes back
A system at equilibrium **shifts to undo** whatever stress you add.

Solubility product
For $AgCl(s) \rightleftharpoons Ag^+ + Cl^-$; the solid is left out.

$$K_{sp} = [Ag^+][Cl^-]$$

EX $K_{sp} = 1.8 \times 10^{-10} \rightarrow s = \sqrt{(1.8 \times 10^{-10})} =$ **1.3×10^{-5} M**

15 ACIDS, BASES, PH & BUFFERS

Three definitions

Theory	Acid	Base
Arrhenius	makes H^+ in water	makes OH^- in water
Brønsted–Lowry	H^+ donor	H^+ acceptor
Lewis	e^--pair acceptor	e^--pair donor

Strong acids HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$ (ionize 100 %)

Strong bases Group 1 hydroxides; $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

MEMORY TRICK

Three halogens, three oxyacids

The six strong acids: **H + Cl, Br, I** and the oxyacids **nitric, sulfuric, perchloric**. Everything else is weak.

Conjugate pairs WORKED

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

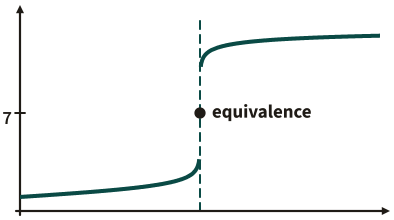
base NH_3 / conjugate acid NH_4^+; acid H_2O / conjugate base OH^-

pH

$[H^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C. pH < 7 acidic, 7 neutral, > 7 basic.

$$pH = -\log[H^+]$$

$$pH + pOH = 14.00$$



25.0 mL of 0.100 M HCl titrated with 0.100 M NaOH: steep jump at 25.0 mL, pH 7 at equivalence.

pH practice WORKED

0.0025 M HCl: pH = −log(0.0025) = **2.60**

0.010 M NaOH: pOH = 2.00 → **pH 12.00**

pH 4.50 → $[H^+] = 10^{-4.50}$ = **3.2×10^{-5} M**

Weak acid: 0.10 M acetic acid WORKED

$K_a = 1.8 \times 10^{-5} = \dfrac{x^2}{0.10 - x}$

$x \ll 0.10$: $\dfrac{x^2}{0.10} \approx 1.8 \times 10^{-5} \rightarrow x = 1.3 \times 10^{-3}$ M

pH = 2.87

Buffers

A weak acid with its conjugate base resists pH change: it soaks up added H^+ or OH^-.

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

EX 0.20 M acetate + 0.10 M acetic acid: 4.74 + log 2 = **pH 5.04**

16 REDOX, CELLS & NUCLEAR CHANGE

Assigning oxidation numbers

Rule	Oxidation number
element alone (Na, O_2)	0
monatomic ion	its charge
O in compounds	−2 (peroxides −1)
H	+1 (with metals −1)
F	always −1
sum	0, or the ion's charge

Find the oxidation number WORKED

$Cr_2O_7^{2-}$: 2x + 7(−2) = −2 → **Cr = +6**

MnO_4^-: x + 4(−2) = −1 → **Mn = +7**

MEMORY TRICK **OIL RIG**

Oxidation **I**s **L**oss of electrons; **R**eduction **I**s **G**ain. (Traditional.)

Balancing in acid (half-reactions)

- Split into oxidation and reduction halves.
- Balance atoms other than O and H.
- O with H_2O; H with H^+; charge with e^-.
- Match electrons, add, cancel.

Permanganate and iron(II) WORKED

$MnO_4^- + 8\,H^+ + 5\,e^- \rightarrow Mn^{2+} + 4\,H_2O$

$5\,Fe^{2+} \rightarrow 5\,Fe^{3+} + 5\,e^-$

$MnO_4^- + 8\,H^+ + 5\,Fe^{2+} \rightarrow Mn^{2+} + 5\,Fe^{3+} + 4\,H_2O$

atoms and charge balance

Standard reduction potentials

Half-reaction	E° (V)
$F_2 + 2e^- \rightarrow 2F^-$	+2.87
$Ag^+ + e^- \rightarrow Ag$	+0.80
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
$2H^+ + 2e^- \rightarrow H_2$	0.00
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$Li^+ + e^- \rightarrow Li$	−3.04

Galvanic (voltaic) cell

- **An**ode: **ox**idation; **red**uction at the **cat**hode (traditional "An Ox, Red Cat").
- Electrons flow anode → cathode through the wire; a salt bridge closes the circuit.
- $E°_{cell} = E°_{cathode} - E°_{anode} > 0$ means spontaneous.

Cell notation

anode | anode ion ‖ cathode ion | cathode (‖ = salt bridge)

Zn | Zn^{2+} ‖ Cu^{2+} | Cu WORKED

E° = 0.34 − (−0.76) = **1.10 V**

ΔG° = −nFE° = −2(96 485)(1.10) = **−212 kJ**

F = 96 485 C/mol

Electroplating copper WORKED

2.00 A for 30.0 min: q = It = 3600 C

3600 ÷ 96 485 = 0.0373 mol e^- → 0.0187 mol Cu = **1.19 g**

$Cu^{2+} + 2e^-$

Radioactive decay

Decay	Particle	Change	Stopped by
alpha	^4_2He	Z − 2, A − 4	paper, skin
beta	$^0_{-1}e$	Z + 1	aluminum
positron	$^0_{+1}e$	Z − 1	aluminum
gamma	$^0_0\gamma$	none (energy)	lead, concrete

MEMORY TRICK

Alpha drops 4 and 2; beta bumps Z

An **alpha** takes 4 off the mass and 2 off Z. A **beta** turns a neutron into a proton, so Z goes **up** one and A stays.

Balance the nuclear equation WORKED

$^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^4_2He$ *238 = 234 + 4; 92 = 90 + 2*

$^{14}_6C \rightarrow {}^{14}_7N + {}^0_{-1}e$

Half-life

$$N = N_0(½)^{t\,/\,t½}$$

Carbon-14 dating, t ½ = 5730 yr WORKED

25.0 % of the original ^{14}C left = 2 half-lives

age ≈ **11 460 yr**

Fission a heavy nucleus splits (U-235 in power plants)

Fusion light nuclei join (the sun's hydrogen → helium)

Mass–energy $E = mc^2$: 1.00 g → 8.99×10^{13} J

Uses medical imaging and treatment, smoke detectors, dating, power