

matter classification

all matter

- physical separation? YES → mixture
 - uniform? YES → homogeneous
 - uniform? NO → heterogeneous
- physical separation? NO → substance
 - chemical decompose? YES → compound
 - chemical decompose? NO → element

stoichiometry strategies

combustion analysis key

- all C → CO₂, all H → H₂O
- O by difference:
 $m_O = m_{\text{sample}} - m_C - m_H$
- trap:** find n_C not n_{C_2}

simultaneous reactions

setup: let x = moles in main. if total moles n , then $(n - x)$ = moles in side rxn

- total reactant: n = (calculate from given mass)
- products equation: mass of products from x and $(n - x)$
- solve for x

LR with yield

- work **backwards** from desired product
- adjust for yield: $n_{\text{needed}} = n_{\text{want}} / \text{yield}$
- don't multiply limiting reagent by $(1 + \text{excess}\%)$. sometimes this is a trap info!

finding empirical formula

- find the moles of each element (**trap:** find n_X not m_{X_2})
- find the simplest whole number ratio

sequential reactions

- find LR in **first** reaction
- product from step 1 becomes reactant in step 2
- check if new reagent limits in step 2

KMT postulates

- gas = very large number of molecules in ceaseless random straight-line motion
- molecules separated by large distance relative to size
- perfectly elastic collisions (no energy loss)
- no intermolecular forces between molecules

temperature measures average translational KE

gas law strategies

chamber mixing

before: P_1V_1 and P_2V_2 separate
after:

$$n_{\text{total}} = \frac{P_1V_1}{RT} + \frac{P_2V_2}{RT} = \frac{P_fV_f}{RT}$$

combustion with T change

- find initial moles: $n_i = \frac{P_iV_i}{RT_i}$
- calculate moles after reaction (balance equation!)
- account for excess reagent remaining
- $P_f = P_i \cdot \frac{n_f}{n_i} \cdot \frac{T_f}{T_i}$ (constant V)

key: excess O₂ leftover contributes to final pressure

piston problems

- piston controls the **pressure**
- free moving piston → constant P

determine ideal gas behavior

- calculate P using $PV = nRT$
- compare to actual P
- if actual $P <$ ideal P , then **attractive forces** dominate
- if actual $P >$ ideal P , then **repulsive forces** dominate
- if equal, then ideal behavior

decomposition in fixed volume

- moles of solid reactant
- count **gas** products only (exclude condensed liquids/solids)
- $P = \frac{nRT}{V}$

percent yield from pressure drop

trick: let x = extent of reaction (mol reacted)

- use the information in initial and final to solve for x
- find out x_{max}
- yield = $\frac{x}{x_{\text{max}}} \times 100\%$

tricky concepts

van der waals adjustments

- high a = strong intermolecular forces
- high b = large molecular volume
- deviations largest at high P , low T

real vs ideal gas

- real → ideal as **density decreases**
- high P , low T = more real (molecules close, interactions matter)
- low P , high T = more ideal (molecules far, move fast)
- ideal gas assumes **no intermolecular forces** and **no molecular volume**

molecular speed distributions

- never say "every molecule of A is faster than B"
- $u \propto \sqrt{T/M}$: lighter/hotter → faster average and broader range
- $u_{\text{rms}} > u_{\text{avg}} > u_{\text{mp}}$ always
- width increases** with temperature (broader range)
- same peak position: higher M needs higher T

liquid vs gas pressure

- gas:** pressure from molecular collisions with walls
- liquid:** pressure from gravitational weight ($P = \rho gh$)

partial pressure

$$P_i = \chi_i P_{\text{total}} \quad \text{where} \quad \chi_i = \frac{n_i}{n_{\text{total}}}$$

ideal gas: mole fraction = volume fraction = pressure fraction

intermolecular forces

- LDF: ALL molecules, **instantaneous induced dipoles**, increases with molar mass
- dipole-dipole: polar molecules only, **permanent dipoles**
- hydrogen bonding: H with FON
- covalent bonding >> H.B. > dipole-dipole > LDF

polarizability

= tendency of electron cloud to distort

- larger molecules → more polarizable
- more electrons → more polarizable
- electrons farther from nucleus → more polarizable
- result:** stronger LDF forces

repulsive forces

- arise from electron orbital overlap
- very strong at short distances
- prevent molecules from collapsing into each other

useful formulas

solution concentration conversions

- % (w/w): g solute per 100 g solution
- % (v/v): mL solute per 100 mL solution
- % (w/v): g solute per 100 mL solution
- ppm: g solute per 10^6 g solution
- ppb: g solute per 10^9 g solution

% excess

$$\% \text{ excess} = \frac{n_{\text{actual}} - n_{\text{stoich}}}{n_{\text{stoich}}} \times 100\%$$

basic gas laws

boyle's law: $P_1V_1 = P_2V_2$ (constant n, T)

charles's law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (constant n, P)

gay-lussac's law: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ (constant n, V)

avogadro's law: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$ (constant P, T)

conversion in ΔT

$$\frac{1 \text{ K}}{1.8^\circ\text{R}} \quad \frac{1 \text{ K}}{1.8^\circ\text{F}}$$

${}^\circ\text{C}$ is the same as K here

dilution

$$M_1V_1 = M_2V_2$$

empirical → molecular

$$n = \frac{M_{\text{actual}}}{M_{\text{empirical}}}$$

then multiply subscripts by n

average molar mass

$$\bar{M} = \sum_i x_i M_i$$

given mole fractions

gas density

$$\rho = \frac{PM}{RT}$$

where M is molar mass

mercury pressure

- $P_{\text{gas}} = P_{\text{bar}} \pm P_{\text{Hg}}$
- observe the diagram!

- $\Delta P_{\text{mmHg}} = \Delta h_{\text{Hg}}$
- watch the units

- just compare the mercury level, no more BS

common traps

- gas products only for $PV = nRT$ (exclude liquids/solids)
- units: K not ${}^\circ\text{C}$, L not mL, atm for $R = 0.082057$
- sig figs rules: $+-$ least decimal places, $\times \div$ least sig figs
- NEVER combine simultaneous rxns into a single equation