

rates of reaction

collision theory

conditions for rxn: collide + sufficient E_a + correct orientation
catalyst:

- increases rate via **lower** E_a path
- participates but **not altered** by rxn
- does **not appear** in overall rxn
- does **NOT** change equilibrium

factors affecting rxn rate:

- concentration of reactants
- temperature
- presence of **catalysts**
- physical nature of reactants

E_p diagrams: enthalpy of reactants and products are **FIXED** regardless of catalyst presence.

reaction rates

must be determined **experimentally**
rxn rate can get the rate of change of single species using **mole ratio**

$$\dot{R} = -\frac{\dot{R}_A}{a} \implies \dot{R}_A = -a \times \dot{R}$$

unit: $M/s (\frac{mol}{L \cdot s})$

positive 4 products, negative 4 reactants

rate laws

relates rate to **concentration** of reactants

differential

for $aA + bB \rightarrow$ products:

$$\dot{R} = -\frac{1}{a} \frac{d[A]}{dt} = -\frac{1}{b} \frac{d[B]}{dt} = k[A]^m[B]^n$$

- m : order w.r.t. A ; n : order w.r.t. B
- $m+n$: overall order
- k : rate constant (only changes with T)

note: exponents \neq coefficients

integrated & half-life

for $aA \rightarrow$ products:

zero-order:

- $[A] = -akt + [A]_0$
- unit of k : M/s
- $t_{1/2} = \frac{[A]_0}{2ak}$

first-order:

- $\ln[A] = -akt + \ln[A]_0$
- unit of k : $1/s$
- $t_{1/2} = \frac{\ln 2}{ak}$

second-order:

- $\frac{1}{[A]} = akt + \frac{1}{[A]_0}$

- unit of k : $1/(M \cdot s)$

$$\bullet t_{1/2} = \frac{1}{ak[A]_0}$$

n-order ($n \neq 1$):

- $\frac{1}{[A]^{n-1}} = (n-1)akt + \frac{1}{[A]_0^{n-1}}$
- unit of k : $\frac{1}{M^{n-1} \cdot s}$
- $t_{1/2} = \frac{2^{n-1}-1}{ak(n-1)[A]_0^{n-1}}$

arrhenius equation

two-point form:

$$\ln \frac{k_2}{k_1} = -\frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

- R is $(8.314 \text{ J}/(\text{mol} \cdot \text{K}))$

phase equilibrium

phase transitions & vapour pressure

phase: region where state of aggregation and chemical composition are uniform. **note:** phase \neq state of matter

phase transition: matter changes from one phase to another due to temperature/pressure changes

vaporization: molecules pass from liquid surface to vapour state

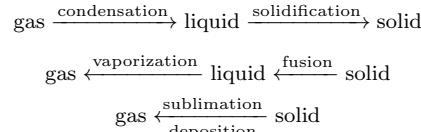
condensation: molecules pass from vapour to liquid state

enthalpy of vaporization:

$$\Delta H_{\text{vap}} = H_{\text{vapour}} - H_{\text{liquid}} = -\Delta H_{\text{condensation}}$$

stronger intermolecular forces \rightarrow higher ΔH_{vap}

phase change relationships:



$$\Delta H_{\text{fus}} + \Delta H_{\text{vap}} = \Delta H_{\text{sub}}$$

phase equilibrium: no net conversion between phases. at equilibrium: vaporization rate = condensation rate

vapour pressure: pressure at equilibrium. weak intermolecular forces \rightarrow high vapour pressure (volatile). strong intermolecular forces \rightarrow low vapour pressure (non-volatile)

$$P_{\text{vap}} = f(T, \text{type of liquid})$$

evaporation vs. boiling:

- evaporation: at surface, any temperature
- boiling: throughout liquid, specific temperature where $P_{\text{vap}} = P_{\text{atm}}$

boiling point: temperature where $P_{\text{vap}}(T_{\text{bp}}) = P$. normal boiling point: $P_{\text{vap}} = 1 \text{ atm}$

higher molecular mass \rightarrow higher polarizability \rightarrow stronger IMF \rightarrow higher bp. exceptions: H_2O , HF , NH_3 (hydrogen bonding)

clausius-clapeyron equation

$$\ln P = -A \left(\frac{1}{T} \right) + B$$

where $A = \frac{\Delta H_{\text{vap}}}{R}$, $R = 8.314 \text{ J}/(\text{mol} \cdot \text{K})$

two-point form:

$$\ln \left(\frac{P_2^{\text{vap}}}{P_1^{\text{vap}}} \right) = -\frac{\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

plot $\ln P_{\text{vap}}$ vs. $\frac{1}{T}$: slope = $-\frac{\Delta H_{\text{vap}}}{R}$

relative humidity:

$$\%RH = \frac{P_{\text{vap}}}{P_{\text{vap}}(T)} \times 100\%$$

dew point: temperature where humid air reaches saturation

phase diagrams

phase diagram: shows regions where phases are in equilibrium (T vs. P)

critical point: no phase boundary between liquid/vapour, $\Delta H_{\text{vap}} = 0$, above critical T only gas exists

supercritical fluid: properties of both gas and liquid (e.g., supercritical CO_2 for decaffeinated coffee)

triple point: solid, liquid, gas coexist

henry's law & raoult's law

solution: homogeneous mixture. solvent: largest amount. solute: smaller amount

concentrations:

- mole fraction: $x_i = \frac{n_i}{\sum n_j}$
- molarity: $M = \frac{\text{moles solute}}{\text{L solution}}$ (temp dependent)
- molality: $m = \frac{\text{moles solute}}{\text{kg solvent}}$

ideal solution: $\Delta H_{\text{soln}} = 0$, similar forces between all components

nonideal solution: $\Delta H_{\text{soln}} \neq 0$

henry's law: gas solubility increases with pressure

$$P_A = H_A x_A \quad \text{or} \quad C_A = k_A P_A$$

where H_A or k_A = henry's constant (depends on solute/solvent pair and temperature)

raoult's law: solute lowers vapour pressure of solvent

$$P_A = x_A P_A^{\text{vap}}$$

for binary ideal solution:

$$P_{\text{total}} = x_A P_A^{\text{vap}} + x_B P_B^{\text{vap}}$$

applies to ideal solutions or dilute solutions ($x_{\text{solv}} > 0.98$)

deviations:

- positive: $P_{\text{total}} > P_{\text{ideal}}$ (weaker unlike interactions)
- negative: $P_{\text{total}} < P_{\text{ideal}}$ (stronger unlike interactions)

colligative properties

depend on ratio of solute particles to solvent molecules: (1) vapour pressure lowering, (2) boiling point elevation, (3) freezing point depression, (4) osmotic pressure

$$\Delta P_{\text{solvent}} = -x_{\text{solute}} P_{\text{solvent}}^{\text{vap}}$$

$$\Delta T_{\text{bp}} = i K_b m \quad \Delta T_{\text{fp}} = -i K_f m$$

van't hoff factor i : particles per formula unit. min: $i = 1$; max: number of ions ($\text{NaCl} \rightarrow 2$, $\text{Pb}(\text{NO}_3)_2 \rightarrow 3$). $i \rightarrow \text{max}$ as dilutes

chemical equilibrium

equilibrium constants

chemical equilibrium: at equilibrium, forward rate = reverse rate
for $aA + bB \rightleftharpoons cC + dD$:
equilibrium constant K_c :

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

where $[]$ = equilibrium concentration in mol/L.
 K_c is constant at given temperature, independent of initial concentrations

equilibrium constant K_P :

$$K_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

partial pressures at equilibrium, use units of bar. K_P is constant at given temperature

relationship:

$$K_P = K_c (RT)^{\Delta n} \quad \text{where } \Delta n = (c+d) - (a+b)$$

$$R = 0.08314 \text{ L}\cdot\text{bar}/(\text{mol}\cdot\text{K})$$

heterogeneous reactions: pure liquids and solids are **not included** in equilibrium constant expressions

magnitude of K :

- $K > 10^{10}$: reaction goes to completion
- $K < 10^{-10}$: reaction does not occur forward
- $10^{-10} < K < 10^{10}$: equilibrium calculation needed

properties of equilibrium constants

- multiply reaction by constant \rightarrow raise K to that power
- reverse reaction: $K_{\text{reverse}} = \frac{1}{K_{\text{forward}}}$
- add reactions: multiply equilibrium constants

reaction quotient & direction

reaction quotient Q :

$$Q_C = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad Q_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

calculated same as K but system doesn't need to be at equilibrium

direction of change:

- $Q_c < K_c$: reactants in excess, forward reaction
- $Q_c = K_c$: at equilibrium
- $Q_c > K_c$: products in excess, reverse reaction

le chatelier's principle

system responds to change by attaining new equilibrium that partially offsets the change
change in concentration:

- add reactant: $Q_c < K_c$, forward reaction, more products
- add product: $Q_c > K_c$, reverse reaction, more reactants

change in volume: reduce volume \rightarrow shifts towards side with fewer gas molecules. effect negligible for condensed phases

change in pressure:

- add/remove gas: affects equilibrium
- change volume: affects equilibrium if gas present
- add inert gas at constant volume: no effect (partial pressures unchanged)

change in temperature:

- endothermic ($\Delta H > 0$): $T \uparrow \rightarrow K \uparrow \rightarrow$ shifts to products
- exothermic ($\Delta H < 0$): $T \uparrow \rightarrow K \downarrow \rightarrow$ shifts to reactants
- reaction shifts away from "energy term" as temperature increases

van't hoff equation:

$$\ln \frac{K_2}{K_1} = -\frac{\Delta H}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

estimate K at different temperature.
clausius-clapeyron is special case where $K =$ vapour pressure

electrochemistry

oxidation states & redox reactions

oxidation state (OS): charge atom would have if all bonds were ionic. not measurable, conceptual tool

rules for assigning OS:

- free element: OS = 0
- monatomic ion: OS = charge
- sum of OS in neutral species = 0
- sum of OS in ion = charge
- H, group 1: +1; group 2: +2; halogens: -1; group 16: -2; group 15: -3

exceptions: H is -1 in hydrides (e.g., LiH); O is -1 in peroxides (e.g., H_2O_2)

redox reactions:

- reduction: OS decreases, gains electrons. oxidizing agent causes oxidation
- oxidation: OS increases, loses electrons. reducing agent causes reduction

balancing redox:

1. write separate half-reactions
2. balance atoms except H, O
3. balance O with H_2O
4. balance H with H^+
5. balance charges with e^-
6. add half-reactions to cancel e^-
7. for basic: add OH^- equal to H^+ to each side, simplify

galvanic cell & faraday's law

galvanic cell: derives electrical energy from spontaneous redox reactions

notation: $Zn(s)|Zn^{2+}(aq)||Cu^{2+}(aq)|Cu(s)$

anode (oxidation): lower potential, electrons flow from anode. **cathode (reduction):** higher potential, electrons flow to cathode

current & faraday's law:

$$I = \frac{Q}{t} \quad \text{units: A = C/s}$$

$$Q = nF \quad n = \frac{Q}{F} = \frac{It}{F}$$

where $F = 96,485 \text{ C/mol}$ (charge of 1 mole of electrons), n = moles of electrons

applications: from moles of electrons, determine moles of substance produced/consumed at electrode

current efficiency:

$$\eta = \frac{\text{charge used}}{\text{charge supplied}} \times 100\% = \frac{\text{actual mass}}{\text{theoretical mass}} \times 100\%$$

standard cell potential & nernst equation

cell potential E_{cell} : electrical potential difference (voltage) between electrodes in V (J/C)

standard state: 1.0 M for dissolved species, 1 bar for gases, 25°C

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

standard: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ with $E^\circ = 0 \text{ V}$

spontaneous direction: electrons travel from low to high potential. higher E° gets reduced (cathode). $E_{\text{cell}}^\circ > 0$ for spontaneous

nernst equation:

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln Q$$

where $R = 8.314 \text{ J/(mol}\cdot\text{K)}$, $T =$ temperature (K), $F = 96,485 \text{ C/mol}$, $Q =$ reaction quotient (aq: M, gas: bar), $n =$ electrons transferred at 25°C:

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0592}{n} \log Q$$

concentration cell: same material at anode and cathode. higher concentration acts as cathode

nernst at equilibrium & electrolytic cells

at equilibrium: $E_{\text{cell}} = 0, Q = K$

$$0 = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln K$$

$$K = \exp \left(\frac{E_{\text{cell}}^\circ nF}{RT} \right)$$

use to determine equilibrium constants with galvanic cells

galvanic vs. electrolytic:

- galvanic: derives electrical energy from spontaneous redox ($E_{\text{cell}} > 0$)
- electrolytic: uses electrical energy to promote non-spontaneous reaction ($E_{\text{cell}} < 0$)

electrochemical cells:

- anode: always oxidation. galvanic: (-); electrolytic: (+)
- cathode: always reduction. galvanic: (+); electrolytic: (-)