

Chemistry of solutions equations (NYB)

Properties of Solutions	
<p>Concentrations</p> <ul style="list-style-type: none"> • The solute (called A): is the compound dissolved in the solvent. • The solvent: is the compound dissolving the solute A • The solution: is the combination of both A and the solvent. 	$\text{molarity (M)} = \frac{\text{mole of solute}}{\text{liter of solution}} = \text{mol/L}$
	$\text{normality (N)} = \frac{\text{equivalent of solute}}{\text{liter of solution}} = \text{Eq/L} \quad (\text{mostly used for acid-base or redox reaction})$ <p>Normality could be ambiguous. The equivalence unit (Eq) depends both on the reactant and the type of reaction present. If H₂SO₄ is used in an acid-base reaction (neutralization); then 1 mole of H₂SO₄ = 2 equivalents of H⁺. Therefore, 1 M H₂SO₄ = 2 N H⁺ (correctly, it should be written H₃O⁺). However, if H₂SO₄ is used in a precipitation reaction with Ba²⁺ then 1 mol H₂SO₄ reacts with 1 mol Ba²⁺ then 1 M H₂SO₄ = 1 N H₂SO₄ IUPAC and NIST discourage the use of normality.</p>
	$\text{molality (m)} = \frac{\text{mole of solute}}{\text{kg of solvent}} = \text{mol/kg}$
	$\text{mass percent (A\%)} = \frac{\text{mass of A}}{\text{total mass of solution}} \times 100\% = \%$
	$\text{mole fraction } (\chi_A) = \frac{\text{mole of A}}{\text{total mole of the solution}} \quad \text{no unit.}$
	$\% \text{dissociation} = \frac{\text{mole of solute ionized}}{\text{mole of formula unit added}} \times 100\% = \%$
Solubility of gases	
Henry's law	$C_{\text{gas}} = k_{\text{gas}} P_{\text{gas}}$ <p>where C_{gas} : gas solubility (mol/L), k_{gas} : Henry's constant for the solubility of the gas (M/atm)</p>
Dalton's law of partial pressure	$P_{\text{total}} = P_A + P_B + P_C + \dots + P_n$ $\chi_A(\text{gas phase}) = \frac{P_A}{P_{\text{Tot}}}$
Colligative properties	
Raoult's law (vapor pressure of two volatile chemicals)	$P_{\text{total}} = \chi_A \cdot P^\circ_A + \chi_B \cdot P^\circ_B$
Raoult's law (solvent + non volatile solute)	$P_{\text{solution}} = \chi_{\text{solvent}} \cdot P^\circ_{\text{solvent}}$
Boiling point elevation	$\Delta T_b = K_b \cdot m_{\text{solute}} \quad (\text{for an electrolyte: } \Delta T_b = i K_b \cdot m_{\text{solute}})$
Freezing point depression	$\Delta T_f = K_f \cdot m_{\text{solute}} \quad (\text{for an electrolyte: } \Delta T_f = i K_f \cdot m_{\text{solute}})$
Osmotic pressure	$\Pi = M R T \quad (\text{for an electrolyte: } \Pi = i M R T)$
van't Hoff factor	$i = \frac{\text{mole of particle in solution}}{\text{mole of formula unit added}} \quad \text{no unit.}$

Equilibrium	
Equilibrium constant • For the equilibrium: $aA + bB \rightleftharpoons cC + dD$	$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$ note: $K = K_c$ (based on molarity)
Reaction quotient (Q)	$Q = \frac{[C]_o^c [D]_o^d}{[A]_o^a [B]_o^b}$ note: $[]_o$ = initial concentration (not at equilibrium)
Relation between K_c and K_p • K_p : pressures in atmosphere • Δn : difference between $n_{\text{gas product}} - n_{\text{gas reactant}}$ • R : $0.08206 \text{ L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}$	$K_p = K_c(RT)^{\Delta n}$
Acid-base and pH	
acid strength	$\% \text{ dissociation} = \frac{\text{number of mole dissociated at equilibrium}}{\text{number of mole initially added}} \times 100\%$
pH measurement	$\text{pH} = -\log[\text{H}^+] \quad , \quad [\text{H}^+] = 10^{-\text{pH}}$ $\text{pOH} = -\log[\text{OH}^-] \quad , \quad [\text{OH}^-] = 10^{-\text{pOH}}$ $\text{pH} + \text{pOH} = 14.00 \text{ at } 25^\circ\text{C}$ $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$
Henderson - Hasselbach equation	$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$ for the conjugate system $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$
Solubility	
solubility product	$\text{MX}_3(\text{s}) \rightleftharpoons \text{M}^{3+}(\text{aq}) + 3\text{X}^-(\text{aq}) \quad \text{then} \quad K_{\text{sp}} = [\text{M}^{3+}][\text{X}^-]^3$
Kinetic	
Rate of a reaction • Consider the reaction: $aA \rightarrow bB$	$\text{Rate} = \frac{\text{change of concentration}}{\text{change in time}} = \frac{-\Delta[\text{A}]}{a \Delta\text{time}} = \frac{+\Delta[\text{B}]}{b \Delta\text{time}} = \text{M/s}$
Rate law • k : rate constant (variable units) • n, m : order of the reactant • $n + m$ = overall order of the reaction	$\text{Rate} = k [\text{A}]^n$ single reactant. $\text{Rate} = k [\text{A}]^n [\text{B}]^m \dots$ multiple reactants.

	Order		
	zero	one	two
Rate law	Rate = k	Rate = $k[A]$	Rate = $k[A]^2$
Integrated rate law	$[A] = [A]_0 - kt$	$\ln[A] = \ln[A]_0 - kt$	$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$
Plot axes to get a line	$[A]$ vs. t	$\ln[A]$ vs. t	$\frac{1}{[A]}$ vs. t
Rate constant units	$M.s^{-1}$	s^{-1}	$M^{-1}.s^{-1}$
Half-life ($t_{1/2}$)	$t_{1/2} = \frac{[A]_0}{2k}$	$t_{1/2} = \frac{\ln 2}{k}$	$t_{1/2} = \frac{1}{k[A]_0}$

Arrhenius equation	<ul style="list-style-type: none"> A = frequency factor ($M.s^{-1}$) $R = 8.314 J.K^{-1}.mol^{-1}$ T = temperature in Kelvin E_a = activation energy ($J.mol^{-1}$) $k = A e^{-E_a/RT}$
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Arrhenius equation (linearized form)	$\ln k = \ln A - \frac{E_a}{R} \times \frac{1}{T}$ <p>The plot $\ln k$ of $1/T$ should give a straight line with a slope of $-E_a/R$.</p>
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Arrhenius equation (for two temperatures)	$\ln \frac{k_2}{k_1} = \frac{-E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$
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Thermodynamic

Difference between two states (Δ)	$\Delta = \text{Final state} - \text{Initial state}$
First law of thermodynamic: conservation of the energy	$\Delta E = q + w$ where E : potential energy, q : heat, w : work $\Delta E = \Delta H - P\Delta V$ where H : enthalpy, P : pressure (kPa), V : volume (L)
Entropy	$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surrounding}}$ $S_{\text{solid}} < S_{\text{liquid}} \ll S_{\text{gas}}$ $\Delta S = \frac{q_{\text{reversible}}}{T}$
Second law of thermodynamic: spontaneity	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ where G = free energy ($kJ.mol^{-1}$), T : temperature (K), S : Entropy ($J.K^{-1}.mol^{-1}$)
Available free energy (system not at equilibrium)	$\Delta G = RT \ln(Q/K)$ where $R = 8.314 J.K.mol^{-1}$, T : temperature (K) $\Delta G = \Delta G^\circ + RT \ln Q$ where: $\Delta G^\circ = -R T \ln K$ The symbol degree on ΔG° indicate "standard state conditions"
Standard state conditions	Gas: $P^\circ = 1 \text{ atm}$ Concentration: $[c]^\circ = 1 \text{ mol/L}$ State of matter: The one at $P = 1 \text{ atm}$ and $T = 25^\circ\text{C}$