

UNIT 1		UNIT 1 (cont)		UNIT 2 (cont)		UNIT 3	
Formula Name	Equation	Formula Name/Subject	Equation/Description	L'C: increase pressure	reaction shifts towards fewer moles	Formula Name/Info	Equation/Description
Boyle's Law	$PV = k (N,T)$	Arrhenius Equation	$Ae^{-E_a/RT}$	L'C: decrease pressure	reaction shifts towards greter moles	Molarity (M)	mol solute/L solution (mol/L)
Charles's Law	$V/T = k (N,P)$	Logarithmic Form of Arrhenius Equation	$\ln(k_2/k_1) = (-E_a/R) * (1/T_2 - 1/T_1)$	L'C: increase volume	reaction shifts towards greater moles	Molality (m)	mol solute/kg solution (mol/kg)
Avogadro's Law	$V/n = k (T,P)$	Michaelis-Menton Enzyme Kinetics	rate = $k_2 [E]T / [S]/[S] + k_m$	L'C: decrease volume	reaction shifts towards fewer moles	Mole Fraction (X)	mol solute/mol solution (mol/mol)
Ideal Gas Law	$V = nRT/P$ or $PV = nRT$	UNIT 2		Endothermic (T) Phase Transitions	Melting, Sublimation, Evaporation	Vapor Pressure (two volatile chemicals)	$P_{vap} = P_a X_a + P_b X_b$
Dalton's Law	$P_{total} = P_1 + P_2 + P_3 \dots$	Equilibrium Constant (concentrations)	$K_c = \frac{[products]}{[reactants]}$	Exothermic (P) Phase Transitions	Freezing, Deposition, Condensation	Boiling Point Elevation	$T_b = k_b m_i$
STP	1 atm, 0 Celsius	Equilibrium Constant (pressures)	$K_p = \frac{P(products)}{P(reactants)}$	Entropy	$S = k_B \ln(W)$	Freezing Point Depression	$T_f = -k_f m_i$
Gas Density	$d = m/V = PM/RT = N/V$	Equilibrium Equation (with rates)	$K = k_f/k_r$	Change in Entropy for Change in Volume	$\Delta S = nR \ln(V_2/V_1)$	Osmotic Pressure	$\pi = iMRT$
Partial Pressure	$P_1 = X_1 * P_{total}$	$Q = K$	At equilibrium	Change in Entropy of Surroundings	$\Delta S_{surr} = -\Delta H/T$	Solubility	$S = \text{mol dissolved} / \text{L solution}$
Root Mean Squared Speed	$v = \sqrt{3RT/M}$	$Q < K$	reaction shifts in forward direction	Gibbs Free Energy	$\Delta G = \Delta H - T\Delta S$	CIE: add in species already present	decrease solubility
Collision Frequency	$f = v_{rms} / mfp$ (mean free path)	$Q > K$	reaction shifts in reverse direction	Nonstandard Gibbs Free Energy (1)	$\Delta G = \Delta G_0 + RT \ln Q$	Henderson-Hasselbach Equation	$pH = pK_a + \log(\text{conj. base/acid})$
Average Kinetic Energy	$KE = 3/2 nRT$	Keq is large (10^3)	$k_{forward} > k_{reverse}$	Standard Gibbs Free Energy	$\Delta G_0 = -RT \ln K$	K value for Neutralization	$K_n = K_a * K_b / K_w$
Effusion - Graham's Law	$rate_1/rate_2 = \sqrt{MW_2/MW_1}$	Keq is small (10^{-3})	$k_{forward} < k_{reverse}$	Nonstandard Gibbs Free Energy (2)	$\Delta G = RT \ln(Q/K)$	Equivalence Point	mol base added = mol acid originally present
General Rate Law	rate = $k[A]^m [B]^n$	Keq = 1	$k_{forward} = k_{reverse}$	Clausius-Claperyon Equation	$\ln(P_2/P_1) = -\Delta H/R (1/T_2 - 1/T_1)$	Midpoint	$pH = pK_a$
Zeroth Order Integrated Rate Law	$[A]_t = [A]_0 - kt$	Kc and Kp Relation	$K_c = K_p * RT^{\Delta n}$	L'C: add more reactants	reaction shifts forward	Solvent-Solvent Interactions	endothermic, overcome IMFs
First Order Integrated Rate Law	$\ln[A]_t = \ln[A]_0 - kt$	L'C: add more products	reaction shifts reverse	Solute-Solute Interactions	endothermic, overcome IMFs		
Second Order Integrated Rate Law	$1/[A]_t = 1/[A]_0 + kt$						
Zeroth Order Half Life	$t = [A]_0/2k$						
First Order Half Life	$t = \ln 2/k$						
Second Order Half Life	$t = 1/k[A]_0$						

UNIT 3 (cont)	
Solvent-Solute Interaction	exothermic, release energy through attractive forces
To relate Ka to Kb	$K_w = K_a * K_b$
Strong Acids	H ₂ SO ₄ , HNO ₃ , HCl, HI, HClO ₄ , HClO ₃ , HBr
Strong Bases	LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH) ₂ , Sr(OH) ₂

UNIT 4 (cont)	
Standard Hydrogen Electrode (SHE)	$2H^+ + 2e^- \rightarrow H_2(g)$
Standard Cell Potential	$E^0 = RT/nF \ln K$
Total Charge (Z)	$Z = I * t$ (in seconds), in coulombs
Moles of Electrons Produced	$n(e^-) = Z/F$
Moles of Metal Produced	$n(\text{metal}) = n(e^-) * \text{mol}(\text{metal})/\text{mols e}^- \text{ needed}$
Mass of Metal Produced	$m(\text{metal}) = n(\text{metal})/MW_{\text{metal}}$
Nernst Equation with pH	$E = E^0_{\text{red}} + 2.3RT/F \log Q$

UNIT 4	
Formula Name/Subject	Equation/Description
Current (I)	charge/time (Amperes)
Voltage (V)	energy/charge (Volts)
Power (W)	energy/seconds (Watts)
Oxidation Reaction	Lose electrons
Reduction Reaction	Gain electrons
Cathode	Site of reduction, e-flow to cathode
Anode	Site of oxidation, e-flow from anode
Cell Potential (E)	$E_{\text{cell}} = E_{\text{cathode}} + E_{\text{anode}} = E_{\text{redcat}} - E_{\text{redan}}$
Gibbs Free Energy (standard)	$\Delta G^0 = -nFE^0$
Faraday's Constant	96,500 C/mol e-
Nernst Equation	$E = E^0 - RT/nF \ln Q$

