

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	$PV = k (N,T)$	Equilibrium Constant (concentrations)	$K_c = \frac{[\text{products}]}{[\text{reactants}]}$	L'C: decrease pressure	reaction shifts towards greater moles	Molarity (M)	mol solute/L solution (mol/L)
Charles's Law	$V/T = k (N,P)$	Equilibrium Constant (pressures)	$K_p = \frac{P(\text{products})}{P(\text{reactants})}$	L'C: increase volume	reaction shifts towards greater moles	Molality (m)	mol solute/kg solution (mol/kg)
Avogadro's Law	$V/n = k (T,P)$	Equilibrium Equation (with rates)	$K = k_f/k_r$	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 $PV = nRT$	Q = K	At equilibrium	Endothermic (T) Phase Transitions	Melting, Sublimation, Evaporation	Vapor Pressure (two volatile chemicals)	$P_{\text{vap}} = P_a X_a + P_b X_b$
Dalton's Law	$P_{\text{total}} = P_1 + P_2 + P_3 \dots$	Q < K	reaction shifts in forward direction	Exothermic (P) Phase Transitions	Freezing, Deposition, Condensation	Boiling Point Elevation	$T_b = k_b m_i$
STP	1 atm, 0 Celsius	Q > K	reaction shifts in reverse direction	Entropy	$S = k_B \ln(W)$	Freezing Point Depression	$T_f = -k_f m_i$
Gas Density	$d = m/V = PM/RT = N/V$	Keq is large (10^3)	$k_{\text{forward}} > k_{\text{reverse}}$	Change in Entropy for Change in Volume	$\Delta S = nR \ln(V_2/V_1)$	Osmotic Pressure	$\pi = iMRT$
Partial Pressure	$P_1 = X_1 \cdot P_{\text{total}}$	Keq is small (10^{-3})	$k_{\text{forward}} < k_{\text{reverse}}$	Change in Entropy of Surroundings	$\Delta S_{\text{surr}} = -\Delta H/T$	Solubility	S = mol dissolved / L solution
Root Mean Squared Speed	$v = \sqrt{3RT/M}$	Keq = 1	$k_{\text{forward}} = k_{\text{reverse}}$	Gibbs Free Energy	$\Delta G = \Delta H - T\Delta S$	CIE: add in species already present	decrease solubility
Collision Frequency	$f = v_{\text{rms}} / \text{mfp}$ (mean free path)	Kc and Kp Relation	$K_c = K_p \cdot RT^{\Delta n}$	Nonstandard Gibbs Free Energy (1)	$\Delta G = \Delta G_0 + RT \ln Q$	Henderson-Hasselbach Equation	$\text{pH} = \text{pK}_a + \log(\text{conj. base/acid})$
Average Kinetic Energy	$KE = 3/2 nRT$	L'C: add more reactants	reaction shifts forward	Standard Gibbs Free Energy	$\Delta G_0 = -RT \ln K$	K value for Neutralization	$K_n = K_a \cdot K_b / K_w$
Effusion - Graham's Law	$\text{rate}_1/\text{rate}_2 = \sqrt{MW_2/MW_1}$	L'C: add more products	reaction shifts reverse	Nonstandard Gibbs Free Energy (2)	$\Delta G = RT \ln(Q/K)$	Equivalence Point	mol base added = mol acid originally present
General Rate Law	$\text{rate} = k[A]^m[B]^n$			Clausius-Clapeyron Equation	$\ln(P_2/P_1) = -\Delta H/R(1/T_2 - 1/T_1)$	Midpoint	$\text{pH} = \text{pK}_a$
Zeroth Order Integrated Rate Law	$[A]_t = [A]_0 - kt$					Solvent-Solvent Interactions	endothermic, overcome IMFs
First Order Integrated Rate Law	$\ln[A]_t = \ln[A]_0 - kt$					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)		UNIT 4 (cont)	
Solvent-Solute Interaction	exothermic, release energy through attractive forces	Standard Hydrogen Electrode (SHE)	$2H^+ + 2e^- \rightarrow H_2(g)$
To relate K_a to K_b	$K_w = K_a * K_b$	Standard Cell Potential	$E^0 = RT/nF \ln K$
Strong Acids	$H_2SO_4, HNO_3, HCl, HI, HClO_4, HClO_3, HBr$	Total Charge (Z)	$Z = I * t$ (in seconds), in coulombs
Strong Bases	$LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)_2, Sr(OH)_2$	Moles of Electrons Produced	$n(e^-) = Z/F$
UNIT 4		Moles of Metal Produced	$n(\text{metal}) = n(e^-) * \text{mol}(\text{metal})/\text{mols e}^- \text{ needed}$
Formula Name/Subject	Equation/Description	Mass of Metal Produced	$m(\text{metal}) = n(\text{metal})/MW_{\text{metal}}$
Current (I)	charge/time (Amperes)	Nernst Equation with pH	$E = E^0_{\text{red}} + 2.3RT/F \log Q$
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	$E = E^0 - RT/nF \ln Q$		

