

# DAT General Chemistry Equation Sheet

## Chapter 0: General and Lab Concepts Review

<b>Dilutions</b>	$M_1V_1 = M_2V_2$ or $C_1V_1 = C_2V_2$	$M$ or $C$ = concentration $V$ = volume
<b>Percent Error</b>	$\frac{(A - T)}{T} \times 100$	$T$ = theoretical $A$ = actual
<b>Absorbance (Spectrophotometer)</b>	$Abs = \epsilon cl$	$\epsilon$ = molar extinction coefficient (molar absorptivity) $c$ = sample's concentration $l$ = path length

## Chapter 2: Atomic and Electronic Structure

<b>Energy of a photon</b>	$E_{photon} = hf = \frac{hc}{\lambda}$	$h$ = Planck's constant ( $6.63 \times 10^{-34} J \cdot s$ ) $f$ = photon's frequency $c$ = speed of light ( $3.0 \times 10^8 m/s$ ) $\lambda$ = photon's wavelength
<b>Absorption/Emission Line Spectra</b>	$\Delta E = E_{photon}$	
<b>Kinetic Energy of an electron (Photoelectric Effect)</b>	$KE_{e^-} = E_{photon} - \phi$	$\phi$ = work function (minimum energy needed to ionize electron)

## Chapter 7: Chemical Solutions

<b>Molarity</b>	$M = \frac{moles_{solute}}{L_{solution}}$	
<b>Molality</b>	$m = \frac{moles_{solute}}{kg_{solvent}}$	
<b>Henry's Law</b>	$P_A = k_H[A]$	$P_A$ = partial pressure of gas A $k_H$ = Henry's Law constant (varies per problem) [A] = conc. of gas A
<b>Freezing Point Depression</b>	$\Delta T_F = -iK_F m$	$i$ = van't Hoff factor $K_F$ = F.P. depression constant $m$ = molality
<b>Boiling Point Elevation</b>	$\Delta T_B = iK_B m$	$i$ = van't Hoff factor $K_B$ = B.P. depression constant $m$ = molality
<b>Vapor Pressure Depression (Raoult's Law)</b>	$P_{soln} = \chi_{solv} P_{solv}^0$	$P_{soln}$ = VP of solution $\chi_{solv}$ = mol fract of solvent $P_{solv}^0$ = VP of pure solvent
<b>Osmotic Pressure (<math>\pi</math>)</b>	$\pi = iMRT$	$M$ = molarity of solute $i$ = van't Hoff factor $R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$ $T$ = temp. in Kelvin

## Chapter 5: Gases

<b>Pressure</b>	$P = \frac{F}{A}$	$F$ = force $A$ = area
<b>Average Kinetic Energy</b>	$KE_{avg} = \frac{3}{2} RT$	$R = 8.314 \frac{J}{mol \cdot K}$
<b>Root-Mean-Square Speed (<math>v</math>)</b>	$v = \sqrt{\frac{3RT}{M_m}}$	$R = 8.314 \frac{J}{mol \cdot K}$ $M_m$ = molar mass
<b>Ideal Gas Law</b>	$PV = nRT$	$n$ = # of moles $R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$
<b>Boyle's Law</b>	$V \propto \frac{1}{P}$	
<b>Charles' Law</b>	$V \propto T$	
<b>Avogadro's Law</b>	$V \propto n$	
<b>Combined Gas Law</b>	$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$	
<b>Standard Temp. &amp; Pressure (STP)</b>	$P=1$ atm $T=273$ K	*1 mol of gas = 22.4 L at STP
<b>Standard Conditions</b>	All aqueous species @ 1M All gaseous species @ 1 atm $T=298$ K	
<b>Density</b>	$\frac{P(MM)}{RT} = \frac{m}{v}$	$MM$ = molar mass $R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$ $m$ = mass $v$ = volume
<b>Dalton's Law of Partial Pressures</b>	$P_{total} = P_A + P_B + \dots$	
<b>Dalton's Law of Partial Pressures</b>	$P_A = \chi_A P_{total}$	$\chi_A$ = mol fraction of gas A
<b>Graham's Law of Effusion</b>	$\frac{r_1}{r_2} = \sqrt{\frac{M_{m2}}{M_{m1}}}$	$r$ = rate of effusion $M$ = molar mass
<b>Real Gas Equation</b>	$(P + \frac{an^2}{V^2})(V - nb) = nRT$	$a$ & $b$ = constants specific to each gas $\frac{an^2}{V^2}$ corrects for IMFs $-nb$ corrects for volume

## Chapter 8: Chemical Kinetics

<b>General Rate Law</b>	$A + B \rightarrow C + D$ $rate = k[A]^m[B]^n$	$k$ = rate constant $m$ & $n$ = determined experimentally
<b>Rate Constant Units</b>	0 order: $k = M^1 \cdot s^{-1}$ 1 <sup>st</sup> order: $k = s^{-1}$ 2 <sup>nd</sup> order: $k = M^{-1} \cdot s^{-1}$ 3 <sup>rd</sup> order: $k = M^{-2} \cdot s^{-1}$	$k$ = rate constant $M$ = molarity $s$ = seconds
<b>Arrhenius Equation</b>	$k = Ae^{-E_a/RT}$	$k$ = rate constant $A$ = unique to each rxn $E_a$ = act. energy $R = 8.314 \frac{J}{mol \cdot K}$ $T$ = temp. in Kelvin

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## Chapter 9: Chemical Equilibria

<b>Equilibrium Constant Expressions</b>	$K_c = \frac{[\text{products}]}{[\text{reactants}]}$	$k = \text{rate constant}$ $P = \text{pressure}$
	$K_{eq} = \frac{k_{\text{forward}}}{k_{\text{reverse}}}$	
	$K_p = \frac{P_{\text{products}}}{P_{\text{reactants}}}$	
<b>Reaction Quotient (Q)</b>	$Q = \frac{[\text{products}]}{[\text{reactants}]}$	$Q > K = \text{shift left}$ $Q < K = \text{shift right}$ $Q = K = \text{equilibrium}$
<b>Solubility Product Constant (<math>K_{sp}</math>)</b>	$K_{sp} = \frac{[\text{products}]}{[\text{reactants}]}$	

## Chapter 10: Acid-Base Equilibria & Titrations

<b>Ionization Constant of Water</b>	$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}$ @ 25 °C	
<b>pH &amp; pOH</b>	$pH = -\log[H^+]$ $pOH = -\log[OH^-]$ $pH + pOH = 14$	$[H^+]$ = conc. of protons $[OH^-]$ = conc. of hydroxide
<b>[H<sup>+</sup>] &amp; [OH<sup>-</sup>]</b>	$[H^+] = 10^{-pH}$ $[OH^-] = 10^{-pOH}$ $[H^+][OH^-] = 1 \times 10^{-14}$	
<b>Weak Acids</b>	$HA + H_2O \rightleftharpoons H_3O^+ + A^-$ $K_a = \frac{[H_3O^+][A^-]}{[HA]}$ $[H^+] = \sqrt{K_a[HA]}$	$K_a = \text{acid dissociation constant}$
<b>Weak Bases</b>	$A^- + H_2O \rightleftharpoons HA + OH^-$ $K_b = \frac{[OH^-][HA]}{[A^-]}$ $[OH^-] = \sqrt{K_b[A^-]}$	$K_b = \text{base dissociation constant}$
<b>pK<sub>a</sub> &amp; pK<sub>b</sub></b>	$pK_a = -\log[K_a]$ $pK_b = -\log[K_b]$ $pK_a + pK_b = 14$ $K_w = K_a \times K_b = 1 \times 10^{-14}$	Larger $K_a = \text{smaller } pK_a$ =stronger acid Larger $K_b = \text{smaller } pK_b$ =stronger base
<b>Neutralization Reaction</b>	$n_A M_A V_A = n_B M_B V_B$	$n_A = \# \text{ of moles } H^+$ $n_B = \# \text{ of moles } OH^-$
<b>Buffers</b>	$pH = pK_a + \log \frac{[A^-]}{[HA]}$	$[A^-] = \text{conc. of base}$ $[HA] = \text{conc. of acid}$

## Chapter 13: Nuclear Reactions

<b>Kinetics (always 1<sup>st</sup> order)</b>	$N = N_0 e^{-kt}$ $\ln N = \ln N_0 - kt$ $t_{1/2} = \frac{0.693}{k}$	$N = \text{amt of radioisotope after time } t$ $N_0 = \text{initial amt}$ $k = \text{rate constant}$ $t = \text{time}$ $t_{1/2} = \text{half life}$ *note that $t_{1/2}$ is independent of concn for 1 <sup>st</sup> order rxns
<b>Nuclear Binding Energy</b>	$E = \Delta mc^2$	$m = \text{mass (MUST be in kg)}$ $c = \text{speed of light}$ $(3.0 \times 10^8 \frac{m}{s})$

## Chapter 11: Thermodynamics & Thermochemistry

<b>Enthalpy (H)</b>	$(\Delta H > 0): \text{endothermic}$ $(\Delta H < 0): \text{exothermic}$	
<b>Enthalpy of Formation</b>	$\Delta H_f = \sum n \Delta H_f^{\circ}(\text{product}) - \sum n \Delta H_f^{\circ}(\text{reactants})$ $n = \text{coefficient from balanced rxn}$	
<b>First Law of Thermodynamics</b>	$\Delta E = q + w$	$\Delta E = \text{change in internal energy}$ $q = \text{heat}$ $w = \text{work}$
<b>Pressure-Volume Work</b>	$w = -P\Delta V$	$P = \text{external pressure}$ $\Delta V = \text{change in volume}$
<b>Calorimetry Thermal Energy (q)</b>	$q = -C_{\text{calorimeter}}\Delta T$	$C_{\text{calorimeter}} = \text{specific heat of calorimeter}$ $\Delta T = \text{change in temp.}$
<b>Heat Curves &amp; Thermal Energy (q)</b>	$q = mc\Delta T$ $q = m\Delta H_{\text{fusion}}$ $q = m\Delta H_{\text{vaporization}}$	$+q: \text{heat gained by system}$ $-q: \text{heat lost from system}$ $m = \text{mass}$ $c = \text{specific heat}$
<b>Entropy (S)</b>	$\Delta S = \sum n S_{\text{products}} - \sum n S_{\text{reactants}}$ $S_{\text{gas}} > S_{\text{liquid}} > S_{\text{solid}}$ $S_{\text{aq}} > S_{\text{solid}}$	
<b>Bond Dissociation Energy</b>	$\Delta H = \sum \Delta H_{\text{reactants}} - \sum \Delta H_{\text{products}}$ $= \sum \Delta H_{\text{broken}} - \sum \Delta H_{\text{formed}}$	$\text{making bonds} = \text{exothermic } (-\Delta H)$ $\text{breaking bonds} = \text{endothermic } (+\Delta H)$
<b>Gibb's Free Energy (<math>\Delta G</math>)</b>	$\Delta G^{\circ} = \Delta H^{\circ} - T\Delta S^{\circ}$	$\Delta G^{\circ} = \text{standard conditions}$ $\Delta H^{\circ} = \text{enthalpy}$ $T = \text{temp. in Kelvin}$ $\Delta S^{\circ} = \text{entropy}$
<b>Gibb's Free Energy (<math>\Delta G</math>)</b>	$\Delta G = \Delta G^{\circ} + RT \ln Q$ $\Delta G^{\circ} = -RT \ln K_{eq}$	$\Delta G = \text{nonstandard conditions}$ $R = 8.314 \frac{J}{\text{mol}\cdot K}$ $Q = \text{reaction quotient}$ $K_{eq} = \text{equilibrium constant}$

## Chapter 12: Electrochemistry & Redox Reactions

<b>Standard Cell Potential</b>	$E^{\circ} = E^{\circ}_{\text{reduction}} + E^{\circ}_{\text{oxidation}}$ $E^{\circ} = E^{\circ}_{\text{cathode}} + E^{\circ}_{\text{anode}}$	
<b>Nernst Equation</b>	$E_{\text{cell}} = E^{\circ} - \frac{0.0592}{n} \log Q$	$E_{\text{cell}} = \text{nonstandard cell potential}$ $n = \# \text{ of electrons transferred}$ $Q = \text{reaction quotient}$
<b>Faraday's Law</b>	$\frac{\text{mass of product}}{I * t_s * MW_{\text{pdt}}}$ $\frac{\text{moles of product}}{I * t_s}$	$MW = \text{molec. weight}$ $I = \text{current (Amps)}$ $t_s = \text{time (seconds)}$ $n = \# \text{ of electrons transferred}$ $F = \text{Faraday's constant}$ $(96485 \frac{\text{coulombs}}{\text{mol } e^-})$