G-Chem Formula Sheet

Nuclear and Atomic Chemistry

Avogadro's number: $N_A = 6.02 \times 10^{23}$

N_A amu (u) $= 1$ gram

$1\text{ u} = 1.66 \times 10^{-24}\text{ g} = 1.66 \times 10^{-27}\text{ kg}$

$m_p = 1.0073\text{ u},\ \ m_n = 1.0087\text{ u}$

$Z =$ # protons, $N =$ # neutrons

mass defect: $\Delta m = (Zm_p + Nm_n) - m_{\text{nucleus}}$

nuclear binding energy: $E_B = (\Delta m) \times \frac{931\text{ MeV}}{1\text{ u}}$

$1\text{ eV} = 1.6 \times 10^{-19}\text{ J},\ \ 1\text{ MeV} = 10^6\text{ eV}$

$E_{\text{photon}} = hf = hc/\lambda$

electron energy levels: $E_n = \frac{Z^2}{n^2}(-13.6\text{ eV})$
for any 1-electron atom

Electron Configurations

e^- quantum numbers: $n,\ \ell,\ m_\ell,\ m_s$

$n = 1, 2, \ldots \qquad \ell = 0, 1, \ldots, n-1$

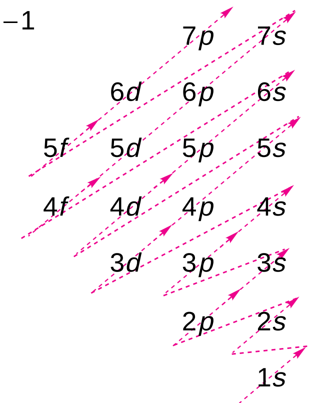
$[\ell = 0 \leftrightarrow s,\ \ell = 1 \leftrightarrow p,$
$\ell = 2 \leftrightarrow d,\ \ell = 3 \leftrightarrow f]$

$m_\ell = -\ell,\ -(\ell-1), \ldots, (\ell-1), \ell$

$m_s = +\frac{1}{2}$ or $-\frac{1}{2}$

in subshell ℓ,
max # of electrons $= 4\ell + 2$

in energy level n,
max # of electrons $= 2n^2$



Radioactive Decay

Z = # protons = atomic number, N = # neutrons,
$A = Z + N$ = mass number

Decay	*Description*	ΔZ	ΔN	ΔA
α	eject $\alpha = {}^4_2\text{He}$	–2	–2	–4
β^-	$n \rightarrow p + e^-$	+1	–1	0
β^+	$p \rightarrow n + e^+$	–1	+1	0
EC	$p + e^- \rightarrow n$	–1	+1	0
γ	$X^* \rightarrow X + \gamma$	0	0	0

Stoichiometry / Lewis Structures

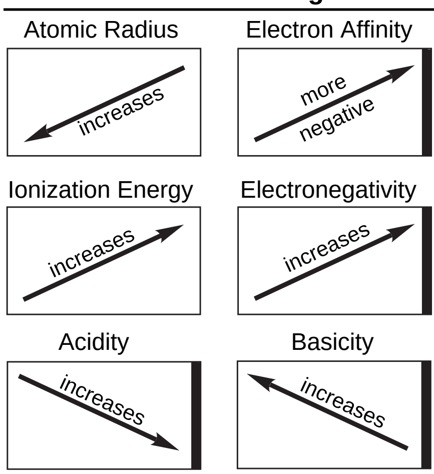
moles $= \frac{\text{mass in grams}}{\text{MW}}$; molarity: $M = \frac{\text{moles of solute}}{\text{L of solution}}$

% composition by mass of X $= \frac{\text{mass of X}}{\text{mass of molecule}} \times 100\%$

formal charge: $FC = V - (\frac{1}{2}B + L)$

$V =$ (# of valence e^-'s), $B =$ (# of bonding e^-'s),
$L =$ (# of lone-pair e^-'s)

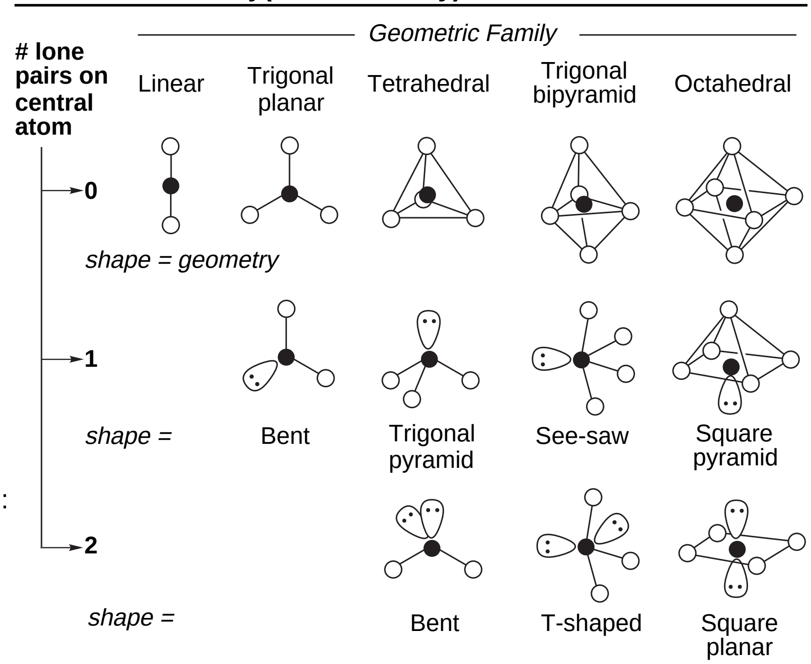
Periodic Trends & Bonding



electronegativity of some common atoms:
F > O > Cl > N > Br > (I ≈ S ≈ C) > H

intermolecular forces (D = dipole,
I = induced, i = instantaneous):
ion–D > D–D (incl. H-bonds)
> D–ID > iD–ID (London)

Molecular Geometry (VSEPR theory)



Gases

Avogadro's law: $V \propto n$

$V_{\text{at STP}} = n(22.4\ \text{L})$

STP: $T = 0\ °\text{C} = 273\ \text{K}$, $P = 1\ \text{atm}$

Boyle's law: $V \propto 1/P$ (at constant T)

Charles' law: $V \propto T$ (at constant P)

Combined: $P_1V_1/T_1 = P_2V_2/T_2$

Ideal-Gas law: $PV = nRT$

Dalton's law of partial pressures: $P = \sum p_i$

Graham's law of effusion:

$$v_{2,\text{rms}} = v_{1,\text{rms}}\sqrt{\frac{m_1}{m_2}} \Rightarrow \frac{\text{rate of effusion of gas 2}}{\text{rate of effusion of gas 1}} = \sqrt{\frac{m_1}{m_2}}$$

Kinetics and Equilibrium

$$\text{concentration rate} = -\frac{\Delta[\text{reactant}]}{\text{time}} \text{ or } +\frac{\Delta[\text{product}]}{\text{time}}$$

$$\text{reaction rate} = -\frac{1}{\text{coeff}}\frac{\Delta[\text{reactant}]}{\text{time}} \text{ or } +\frac{1}{\text{coeff}}\frac{\Delta[\text{product}]}{\text{time}}$$

rate law for *elementary* reaction: $\text{rate} = k[\text{reactant}_1]^{\text{coeff}_1}\cdots$

Arrhenius equation: $k = Ae^{-E_a/RT}$

for generic balanced reaction $a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D}$,

$$\text{equilibrium constant: } K_{eq} = \frac{[\text{C}]^c_{\text{at eq}}[\text{D}]^d_{\text{at eq}}}{[\text{A}]^a_{\text{at eq}}[\text{B}]^b_{\text{at eq}}}$$

excluding pure solids and liquids

(gas rxns use partial pressures in K_{eq} expression)

K_{eq} is a constant at a given temperature.

$K_{eq} < 1 \Leftrightarrow$ equilibrium favors reactants

$K_{eq} > 1 \Leftrightarrow$ equilibrium favors products

$$\text{reaction quotient: } Q = \frac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}$$

Law of Mass Action (Le Châtelier's principle):

$Q < K_{eq} \Leftrightarrow$ rxn proceeds forward

$Q = K_{eq} \Leftrightarrow$ rxn at equilibrium

$Q > K_{eq} \Leftrightarrow$ rxn proceeds in reverse

Acids and Bases

$\text{pH} = -\log[\text{H}^+] = -\log[\text{H}_3\text{O}^+]$

$\text{pOH} = -\log[^-\text{OH}]$

$K_w = [\text{H}^+][^-\text{OH}] = 1 \times 10^{-14}$ at 25 °C

pH + pOH = 14 at 25 °C

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]},\quad \text{p}K_a = -\log K_a$$

$$K_b = \frac{[^-\text{OH}][\text{HB}^+]}{[\text{B}]},\quad \text{p}K_b = -\log K_b$$

$K_aK_b = K_w$ = ion-product constant for water

Henderson–Hasselbalch equations:

$$\text{pH} = \text{p}K_a - \log\frac{[\text{weak acid}]}{[\text{conjugate base}]}$$

$$\text{pOH} = \text{p}K_b - \log\frac{[\text{weak base}]}{[\text{conjugate acid}]}$$

acid–base neutralization: $N_aV_a = N_bV_b$

Colligative Properties

$$\text{molality: } m = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

$$\text{normality: } N = \frac{\text{equivalents (eq)}}{\text{L of solution}}$$

BP elevation: $\Delta T_b = k_b im$

FP depression: $\Delta T_f = -k_f im$

$$\text{mole fraction: } X_S = \frac{\text{moles of S}}{\text{total moles}}$$

Raoult's law: $P_A = X_A P_A^o$

vapor pressure depression: $\Delta P_A = -(1 - X_A)P_A^o$

osmotic pressure: $\Pi = iMRT$

Thermochemistry

T (in K) = $T_{°C} + 273$, 1 cal ≈ 4.2 J

$q = mc\Delta T = C\Delta T$ (no phase change)

$q = n\Delta H_{\text{phase change}}$

enthalpy change: ΔH = heat of rxn at const P

$\Delta H < 0 \Leftrightarrow$ exothermic, $\Delta H > 0 \Leftrightarrow$ endothermic

standard state: one most stable at 25°C, 1 atm

$\Delta H^o_{\text{rxn}} = \sum n\Delta H^o_{\text{f,products}} - \sum n\Delta H^o_{\text{f,reactants}}$

Laws of Thermodynamics:

1) Energy is conserved: $\Delta E = q + W$
2) spontaneous rxn $\Rightarrow \Delta S$ [Δentropy] ≥ 0
3) $S = 0$ for pure crystal at $T = 0$ K

Gibbs Free Energy: $\Delta G = \Delta H - T\Delta S$ [const. T]

$\Delta G < 0 \Leftrightarrow$ spontaneous

$\Delta G = 0 \Leftrightarrow$ at equilibrium

$\Delta G > 0 \Leftrightarrow$ reverse rxn is spontaneous

$$\Delta G^o \approx -RT\ln K \approx -2.3RT\log K \approx (-5.7\ \tfrac{\text{kJ}}{\text{mol}})\log K$$

Redox and Electrochemistry

Rules for determining oxidation state (OS):

1) sum of OS's = 0 in neutral molecule;
   sum of OS's = charge on ion
2) Group 1 metals: $OS = +1$;
   Group 2 metals: $OS = +2$
3) OS of F = –1
4) OS of H = +1
5) OS of O = –2
6) OS of halogens = –1; OS of O family = –2

If one rule contradicts another, rule higher in list takes precedence.

F = faraday ≈ 96,500 C/mol e$^-$

$\Delta G = -nFE_{\text{cell}}$

$E_{\text{cell}} > 0 \Leftrightarrow$ spontaneous

$E_{\text{cell}} < 0 \Leftrightarrow$ reverse rxn is spontaneous

$$\text{Nernst equation: } E = E^o - \frac{0.06}{n}\log Q$$

Faraday's Law of Electrolysis:

The amount of chemical change is proportional to the amount of electricity that flows through the cell.