1 Properties of Solutions

The heat an object gains or loses is proportional to its temperature change.

q = # Moles × Molar heat capacity × ΔT   OR   q = Mass × Spefic Heat Capacity × ΔT

The heat of fusion and vaporization are heats required or given off during phases changes.

q = # Moles × ΔH_{fusion} q = # Moles × $\Delta H_{vaporization}$

When calculating heat changes in matter the conservation of energy must be obeyed.

$\sum q_i = q_{heating\ solid} + q_{melting\ solid} + q_{heating\ liquid} + q_{vaporizing\ liquid} + q_{heating\ gas}$

The Clausius-Clayperon equation relates the vapor pressure P_1 of a solvent at temperature T_1 to the vapor pressure of the solvent P_2 at temperature T_2

$$\ln P = \frac{-\Delta H_{vap}}{R}\left(\frac{1}{T}\right) \qquad \ln \frac{P_2}{P_1} = \frac{-\Delta H_{vap}}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right) \quad \text{where R is 8.314 J mol}^{-1}\text{ K}^{-1}.$$

The solubility of a gas in a liquid is equal to the partial pressure of that gas above the liquid times the Henry's Law constant for that gas/solvent combination.

$S_{gas} = k_{Henry} \times P_{vapor}$

Raoult's Law: The partial pressure of a vapor over its liquid is equal to the mole fraction of the solvent times the vapor pressure of the pure liquid.

$P_v = \chi_{solvent}\ P^{\circ}_{solvent}$

Four colligative properties depend only on the number of solute particles in solution.

Boiling Pt. Elevation	Freezing Pt. Depression	Osmotic Pressure	Vapor Pressure Lowering
$\Delta T_b = (T_b - T_b^{\circ}) = imK_b$	$\Delta T_f = -(T_f - T_f^{\circ}) = imK_f$	$\Pi = iMRT$	$\Delta P_v = i\chi_{solute}\ P^{\circ}_{solvent}$

R = 0.0821 L atm mol^{-1} K^{-1}, m = molalilty, M = molarity, i = Van Hoff't Constant.
In the ideal case i = total number of moles of particles: i.e. i =1 for electrolytes, i = 2 for NaCl, i = 3 for $CaCl_2$ and so on.

Chemistry makes use of many units of concentration.

$$\text{molarity} = M = \frac{\text{moles solute}}{\text{L solution}} \qquad \text{molality} = m = \frac{\text{moles solute}}{\text{kg solvent}} \qquad \%\text{ mass} = \frac{\text{mass solute}}{\text{mass of solution}} \times 100$$

$$\%\text{ volume} = \frac{\text{vol of solute}}{\text{vol of solution}} \times 100 \qquad \text{mole fraction} = \frac{\text{moles of solute}}{\text{total \# of moles}}$$

$$\text{parts per million} = \text{ppm} = \frac{\text{g solute}}{\text{g solution}} \times 10^6 = \frac{\text{g solute}}{\text{mL solution}} \times 10^6 \text{ (when } D_{H2O}\text{: 1.0 g/mL)}$$

mass of solution = mass of solvent + mass of solute

2 Kinetics and Integrated Rate Laws (A $\longleftrightarrow$ C)

| Property	0th Order	1st Order	2nd Order
Rate Law	Rate = k	Rate = k[A]	Rate = k[A]^2
Integrated Rate Law	$[A]_t = -k\,t + [A]_0$	$\ln[A]_t = -k\,t + \ln[A]_0$	$\dfrac{1}{[A]_t} = k\,t + \dfrac{1}{[A]_0}$
Slope, Y-intercept	$-k$, $[A]_0$	$-k$, $\ln[A]_0$	k, $\dfrac{1}{[A_0]}$
Y vs x Plot	$[A]_t$ vs t	$\ln[A]_t$ vs. t	$\dfrac{1}{[A]_t}$ vs t
Half-Life = t$_{1/2}$	$\dfrac{[A]_0}{2k}$	$\dfrac{0.693}{k}$	$\dfrac{1}{k\,[A]_0}$

The Arrhenius equation relates the rate constant to the activation energy, E$_a$ at temperature T.

$$k = A\,e^{-E_a/RT} \qquad \ln\frac{k_2}{k_1} = \frac{-E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right) \qquad R = 8.314 \text{ J mol}^{-1}\text{ K}^{-1}.$$

3 Equilibrium

The law of mass action exclude solids and pure liquids from the equilibrium expression.

$$aA + bB \longrightarrow cC + dD \qquad K_c = \frac{[C]^c[D]^d}{[B]^b[A]^a} \qquad K_p = \frac{(P_C)^c(P_D)^d}{(P_B)^b(P_A)^a}$$
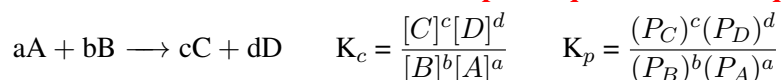
K$_c$ and K$_p$ are related.

$$K_c = K_p(RT)^{-\Delta n} \; (\Delta n = \text{\# mol of products - \# mol of reactants}) \qquad R = 0.0821 \text{ L atm mol}^{-1}\text{ K}^{-1}$$

ICE equiibrium problems simplify when K$_a$ $\times$ 100 < [HA] or when K$_b$ $\times$ 100 < [B]. The quadratic equation must be used when simplications can not be made.
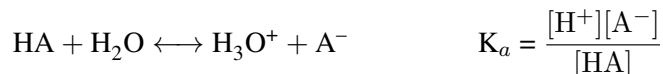
$$ax^2 + bx + c = 0 \qquad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

The van't Hoff equation relates the equilibrium constant, K$_1$ at T$_1$ to the equilibrium constant for the same reaction K$_2$ at T$_2$.

$$\ln\frac{K_2}{K_1} = \frac{-\Delta H^\circ_{rxn}}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right) \qquad R = 8.314 \text{ J mol}^{-1}\text{ K}^{-1}, \text{ T in Kelvin } (\Delta H \text{ must be known}).$$

4 Acids and Bases

K$_a$ (K$_b$) is the equilibrium constant for acids (bases).

$$HA + H_2O \longleftrightarrow H_3O^+ + A^- \qquad K_a = \frac{[H^+][A^-]}{[HA]}$$

$$B + H_2O \longleftrightarrow BH^+ + OH^- \qquad K_b = \frac{[OH^-][BH^+]}{[B]}$$

The Ion-Product of Water is a constant for all aqueous system at 25°C.

$K_w = 1.0 \times 10^{-14}$ = [H$^+$] [OH$^-$] and the log form of the same: pH + pOH = 14

pH and pOH are logarithmic measures of [H$^+$] and [OH$^-$] concentration.

pH = $-\log$ [H$^+$]   pOH = $-\log$ [OH$^-$]   pH + pOH = 14

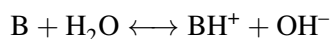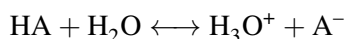**Acid-Conjugate Base Dissociation Constants Are Related through the Ion-Product of Water.**

$K_a \times K_b = K_w = 1.0 \times 10^{-14}$ or $pK_a + pK_b = 14$.

% Ionization of An Acid or a Base Is Found from equilibrium pH data.

$$\% \text{ Ionization} = \frac{[\text{H}^+]_{eq}}{[\text{HI}]_{initial}} \times 100 \qquad \% \text{ Ionization} = \frac{[\text{OH}^-]_{eq}}{[\text{B}]_{initial}} \times 100$$
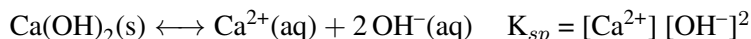
The Henderson-Hasslebalch Equation is used to produce buffers.

$$\text{HA} + \text{H}_2\text{O} \longleftrightarrow \text{H}_3\text{O}^+ + \text{A}^- \qquad\qquad \text{B} + \text{H}_2\text{O} \longleftrightarrow \text{BH}^+ + \text{OH}^-$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]_{init}}{[\text{HA}]_{init}} \qquad\qquad \text{pOH} = \text{p}K_b + \log \frac{[\text{BH}^+]_{init}}{[\text{B}]_{init}}$$

K_{sp} is the equilibrium constant that characterizes the equilibrium between a barely soluble salt and its saturated solution

$$\text{Ca(OH)}_2(\text{s}) \longleftrightarrow \text{Ca}^{2+}(\text{aq}) + 2\,\text{OH}^-(\text{aq}) \quad K_{sp} = [\text{Ca}^{2+}]\,[\text{OH}^-]^2$$

5   Thermodynamics

2nd Law of Thermo: the entropy of the universe increases in spontaneous processes.

$$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings} > 0 \quad \text{equivalent statement is } \Delta G^\circ_{rxn} < 0$$

The standard Gibb's Free Energy, $\Delta \text{G}^\circ_{rxn}$ is used to predict the spontaneity of a reaction.

$\Delta \text{G}^\circ_{rxn} = \Delta \text{H}^\circ_{rxn}$ - T$\Delta \text{S}^\circ_{rxn}$ (T in Kelvin, S and H have the same units of energy).

At standard-state conditions, Hess's Law can be used to calculate thermodynamic quantities.

$\text{S}^\circ_{rxn} = \sum n_i \text{S}^\circ_{if}(\text{products}) - \sum m_j \text{S}^\circ_{jf}(\text{reactants})$

$\Delta \text{H}^\circ_{rxn} = \sum n_i \Delta \text{H}^\circ_{if}(\text{products}) - \sum m_j \Delta \text{H}^\circ_{jf}(\text{reactants})$

$\Delta \text{G}^\circ_{rxn} = \sum n_i \Delta \text{G}^\circ_{if}(\text{products}) - \sum m_j \Delta \text{G}^\circ_{jf}(\text{reactants})$

$\Delta \text{G}^\circ_{rxn} = \Delta \text{H}^\circ_{rxn}$ - T$\Delta \text{S}^\circ_{rxn}$ (T in Kelvin, S and H same units of energy.)

ΔG_{rxn}, can be calculated for reactions under non-standard state conditions.

$\Delta \text{G}_{rxn} = \Delta \text{G}^\circ_{rxn} + RT \ln Q$ (R = 8.314 J mol^{-1} K^{-1} and T in Kelvin, Q is reaction quotient)

At equilibrium: $\Delta \text{G} = 0$ & $\Delta \text{G}^\circ_{rnx}$ is linked to K of a chemical reaction.

$$\Delta \text{G}^\circ_{rxn} = -RT \ln K \qquad K = \exp\left(\frac{-\Delta G^\circ_{rxn}}{RT}\right) \quad (\text{R = 8.314 J mol}^{-1}\text{ K}^{-1} \text{ and T in Kelvin})$$

The cross-over temperature with which a spontaneous may be forced to occur can be calculated.

$$\text{T} > \frac{\Delta H^\circ_{rxn}}{\Delta S^\circ_{rxn}} \text{ (units of energy must match)}$$