

Chem 7/11 Equations In \LaTeX

Stoichiometry & Aqueous Solutions

The mole counts numbers of atoms or molecules using mass.

$$1 \text{ mole} = \text{A.N.} = 6.02 \times 10^{23} \text{ entities} \quad \text{Mol } X = \frac{\text{grams } X}{\text{Molar Mass } X}$$

$$\% \text{ Yield} = \frac{\text{experimental mass}}{\text{theoretical mass}} \times 100$$

% Mass \leftrightarrow Chemical Formula

$$\% \text{ mass of Element } X = \frac{\text{mol } X \text{ in formula} \times \text{molar mass}/1 \text{ mol } X}{\text{mass of 1 mol compound}}$$

Molarity is a common concentration unit in aqueous chemistry. The dilution equation is used for the preparation of solutions using stock solutions.

$$\text{molarity} = M = \frac{\text{moles solute}}{\text{L solution}} \quad M_1V_1 = M_2V_2$$

Gases

The ideal gas law describes the behavior of gases at low pressure and high temperatures.

$$PV = nRT \quad \frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \text{ (move constants to one side variables to another)}$$

The gas law can be expressed as a function of molar mass and gas density.

$$\text{Molar Mass} = \frac{mRT}{VP} = \frac{dRT}{P} \quad 0.0821 \text{ L atm/mol K and T in Kelvin}$$

Dalton's Law: P_{Total} of a mixture of gases is the sum of the partial pressures.

$$P_{total} = P_1 + P_2 + P_3 + \dots$$

Thermochemistry

Internal and energy and enthalpy are state functions

$$\Delta E = q + w = q - P\Delta V \quad \Delta H = q_p = \Delta E + P\Delta V \quad \Delta H_{rxn}^\circ = \Delta H_{products}^\circ - \Delta H_{reactants}^\circ$$

Standard State is a specified experimental conditions of 1 atm ideal gas, 1M liquids and 298K.

Hess' Law and ΔH_f° values allows computation of ΔH_{rxn}° standard state conditions.

$$\sum n_i \Delta H_{if}^\circ(\text{products}) - \sum m_j \Delta H_{jf}^\circ(\text{reactants})$$

Quantum Theory

Energy is quantized at the molecular level. The speed of light in a vacuum is constant.

$$E = h\nu \quad h = 6.63 \times 10^{-34} \text{ J sec} \quad c = \lambda\nu = 3.0 \times 10^8 \text{ m/s}$$

IMF's and Solutions

The heat an object gains or loses is proportional to its mass and the change in temperature.

$$q = \text{Mass} \times \text{Specific Heat Capacity} \times \Delta T \quad \text{OR} \quad q = \# \text{ Moles} \times \text{Molar heat capacity} \times \Delta T$$

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

$$q = \# \text{ Moles} \times \Delta H_{\text{fusion}} \quad \text{and} \quad q = \# \text{ Moles} \times \Delta H_{\text{vaporization}}$$

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

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

The Clausius-Clayperon relates the vapor pressure P_1 at T_1 to the vapor pressure P_2 at T_2 .

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

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.

$$S = k_{\text{Henry}} \times P_i$$

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.

$$P_{\text{vapor}} = \chi_{\text{solvent}} P_{\text{solvent}}^{\circ}$$

Colligative properties are physical properties that depend only on the number of solute molecules dissolved in a solvent (and not the chemical identity of the solute).

BP Elevation	FP Depression	Vapor Pressure Lowering	Osmotic Pressure
$\Delta T_b = (T_b - T_b^{\circ}) = m K_b$	$\Delta T_f = -(T_f - T_f^{\circ}) = m K_f$	$\Delta P_{\text{vapor}} = \chi_{\text{solute}} P_{\text{solvent}}^{\circ}$	$\pi = M R T$

Concentration Units

Chemistry uses many units to specify concentration.

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

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

$$\text{parts per million} = \text{ppm} = \frac{\text{g solute}}{\text{g solution}} \times 10^6 = \frac{\text{g solute}}{\text{mL solution}} \times 10^6 \quad \text{when density}_{H_2O}: 1.0 \text{ g/mL}$$

$$\text{mass of solution} = \text{mass of solvent} + \text{mass of solute}$$

Kinetics

Property	0th Order	1st Order	2nd Order
Rate Law	Rate = k	Rate = k[A]	Rate = k[A] ²
Units of k	mol L ⁻¹ s ⁻¹	s ⁻¹	L mol ⁻¹ s ⁻¹
Integrated Rate Law	[A] _t = -k t + [A] ₀	ln[A] _t = -k t + ln[A] ₀	$\frac{1}{[A]_t} = k t + \frac{1}{[A]_0}$
Y vs x Plot	[A] _t vs t	ln [A] _t vs. t	$\frac{1}{[A]_t}$ vs t
Slope, Y-intercept	-k, [A] ₀	-k, ln [A] ₀	k, $\frac{1}{[A]_0}$
Half-Life = t _{1/2}	$\frac{[A]_0}{2k}$	$\frac{0.693}{k}$	$\frac{1}{k [A]_0}$

1st-Order Decay: N(t) = N₀ exp(-k t)

Integrated Form: ln N(t) = -k t + ln (N₀)

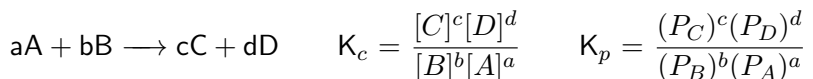
where: Rate Constant = $k = \frac{0.693}{\tau_{1/2}}$ and τ_{1/2} is the half-life

The Arrhenius equation relates the rate constant, k, to the activation energy, E_a at temperature T.

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

Equilibrium

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



The Equilibrium Constants: K_c and K_p Are Related

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

The van't Hoff equation relates K₁ at T₁ to K₂ at T₂

$$\ln \frac{K_2}{K_1} = \frac{-\Delta H_{rxn}^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) \quad R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}, \text{ T in Kelvin.}$$

Acids and Bases

K_a (K_b) is the equilibrium constant for acids (bases).



$$K_a = \frac{[H^+][A^-]}{[HA]} \quad K_b = \frac{[OH^-][BH^+]}{[B]}$$

pH is a logarithmic measure of [H⁺]

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

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

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

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

$$K_a \times K_b = K_w = 1.0 \times 10^{-14} \quad \text{or} \quad \text{p}K_a + \text{p}K_b = 14.$$

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

$$\% \text{ Ionization} = \frac{[\text{H}^+]_{eq}}{[\text{HI}]_{initial}} \times 100 \quad \quad \quad \% \text{ Ionization} = \frac{[\text{OH}^-]_{eq}}{[\text{B}]_{initial}} \times 100$$

ICE equilibrium problems simplify when $K_a \times 100 < [\text{HA}]$ or when $K_b \times 100 < [\text{B}]$. The quadratic equation must be used when simplifications can not be made.

$$ax^2 + bx + c = 0 \quad \quad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

The Henderson-Hasselbalch Equation is used to produce buffers.



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

Thermodynamics

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

$$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings} \quad \quad \quad \Delta G_{rxn}^\circ = \Delta H_{rxn}^\circ - T\Delta S_{rxn}^\circ$$

The standard Gibb's Free Energy, ΔG_{rxn}° is used to predict the spontaneity of a reaction.

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

$$T > \frac{\Delta H_{rxn}^\circ}{\Delta S_{rxn}^\circ} \quad (\text{T in Kelvin, S and H have the same units J or kJ}).$$

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

$$S_{rxn}^\circ = \sum n_i S_{if}^\circ(\text{products}) - \sum m_j S_{jf}^\circ(\text{reactants}) \quad \quad \Delta H_{rxn}^\circ = \sum n_i \Delta H_{if}^\circ(\text{products}) - \sum m_j \Delta H_{jf}^\circ(\text{reactants})$$

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

$$\Delta G_{rxn}^\circ = \Delta H_{rxn}^\circ - T\Delta S_{rxn}^\circ \quad (\text{T in Kelvin, S and H in kJ})$$

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

$$\Delta G_{rxn} = \Delta G_{rxn}^\circ + RT \ln Q \quad R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1} \text{ and T in Kelvin)}$$

ΔG_{rxn}° is linked to the equilibrium constant, K, of a chemical reaction.

$$\Delta G_{rxn}^\circ = -RT \ln K_p \quad R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1} \text{ and T in Kelvin}$$

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