

# Fundamental Chemistry formula booklet

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## Constants and unit conversions

Gas constant	$R = 0.082057 \frac{\text{L atm}}{\text{K mol}}$ $= 8.314 \frac{\text{J}}{\text{K mol}}$
Avogadro constant	$N_A = 6.02 \cdot 10^{23} \text{mol}^{-1}$
Boltzmann constant	$k = 1.38 \cdot 10^{-23} \text{ J/K}$
Planck constant	$h = 6.63 \cdot 10^{-34} \text{ J}\cdot\text{s}$
Rydberg constant	$R_H = 2.18 \cdot 10^{-18} \text{ J}$
Ion-product constant of water	$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \cdot 10^{-14}$
Faraday constant	$F = 9.65 \cdot 10^4 \text{ C/mol(e}^-)$
°C to K	$yK = (x^\circ\text{C} + 273.15) \frac{1K}{1^\circ\text{C}}$
atm to mmHg	$1 \text{ atm} = 760 \text{ mmHg}$
atm to kPa	$1 \text{ atm} = 101.325 \text{ kPa}$

## Stoichiometry (Chapter 3)

Percent composition of an element in a compound	$\text{percent composition} = \frac{n \cdot \text{molar mass of element}}{\text{molar mass of compound}} \cdot 100\%$
% yield	$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \cdot 100\%$

## Reactions in Aqueous Solutions (Chapter 4)

Molarity	$M = \frac{n(\text{solute})}{V(\text{solution})}$
Dilution of solution	$M_i V_i = M_f V_f$

## Gases (Chapter 5)

Ideal gas equation	$\begin{aligned} PV &= nRT \\ \frac{P_1 V_1}{n_1 T_1} &= \frac{P_2 V_2}{n_2 T_2} \end{aligned}$
Ideal gas equation with density and molar mass	$d = \frac{PM}{RT}$
Mole fraction	$X_i = \frac{n_i}{n_T}$
Partial pressure	$P_i = X_i P_T$
Root-mean-square speed of gas molecules	$v_{rms} = \sqrt{\frac{3RT}{M}}$
Rates of diffusion (Graham's law)	$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$
Nonideal gas equation (van der Waals equation)	$\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT$ <p style="text-align: center;"><math>a</math> and <math>b</math> are constants</p>

## Energy and Thermodynamics (Chapter 6 & 18)

Work done by gas expansion or compression	$W = -P\Delta V$
Enthalpy change for a constant-pressure process	$\Delta H = \Delta U + P\Delta V$
Enthalpy change for a constant-temperature process	$\Delta U = \Delta H - RT\Delta n$
Heat capacity	$C = mc$
Heat change from specific heat	$Q = mc\Delta T$
Heat change from heat capacity	$Q = C\Delta T$
Standard enthalpy of reaction	$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{products}) - \sum m\Delta H_f^\circ(\text{reactants})$ <p><math>n</math> and <math>m</math> are the stoichiometric coefficients for the products and reactants</p>
Entropy	$S = k \ln(W)$ <p><math>k</math> is the Boltzmann constant (<math>1.38 \cdot 10^{-23}</math> J/K), <math>W</math> is the number of microstates</p>
Second law of thermodynamics	<p>Spontaneous process: <math>\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr} &gt; 0</math></p> <p>Equilibrium process: <math>\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr} = 0</math></p>
Standard entropy change of reaction	$\Delta S^\circ = \sum n\Delta S^\circ(\text{products}) - \sum m\Delta S^\circ(\text{reactants})$ <p><math>n</math> and <math>m</math> are the stoichiometric coefficients for the products and reactants</p>
Gibbs free energy	$G = H - TS$
Free energy change at constant temperature	$\Delta G = \Delta H - T\Delta S$
Standard free energy change of a reaction	$\Delta G^\circ = \sum n\Delta G_f^\circ(\text{products}) - \sum m\Delta G_f^\circ(\text{reactants})$ <p><math>n</math> and <math>m</math> are the stoichiometric coefficients for the products and reactants</p>

Free energy change from standard free energy change and reaction quotient	$\Delta G = \Delta G^\circ + RT \ln(Q)$ $\Delta G$ is free energy change, $\Delta G^\circ$ is standard free energy change, $R$ is the gas constant ( $8.314 \frac{\text{J}}{\text{K mol}}$ ), $Q$ is the reaction quotient
Standard free energy change from equilibrium constant	$\Delta G^\circ = -RT \ln(K)$

## Electrons and the periodic table (Chapter 7 & 8)

Wave speed	$v = \lambda f$ $\lambda$ is wavelength, $f$ is frequency
Energy of a photon	$E = hf = \frac{hc}{\lambda}$ $h$ is Planck's constant ( $6.63 \cdot 10^{-34} \text{ J}\cdot\text{s}$ )
Photoelectric effect	$hf = E_{kin} + W$
Energy of an electron in the $n$ th state in a hydrogen atom	$E_n = -R_H \left( \frac{1}{n^2} \right)$ $R_H$ is the Rydberg constant ( $2.18 \cdot 10^{-18} \text{ J}$ )
Energy of a photon emitted/absorbed when an electron changes energy level	$\Delta E = hf = R_H \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$
Wavelength in terms of mass and velocity	$\lambda = \frac{h}{mv}$
Effective nuclear charge	$Z_{eff} = Z - \sigma$ $\sigma$ is the shielding constant

## Bonds (Chapter 9 & 10)

Enthalpy change from bond enthalpies	$\Delta H^\circ = \Sigma H_{bond}(\text{reactants}) - \Sigma H_{bond}(\text{products})$
Dipole moment	$\mu = Q \cdot r$ Q is the charge, $r$ is the distance between charges
Bond order	$\text{bond order} = \frac{1}{2} (\text{number of electrons in bonding MOs} - \text{number of electrons in antibonding MOs})$

## Intermolecular interactions and solutions (Chapter 12 & 13)

Molar heat of vaporization (Clausius-Clapeyron equation)	$\ln(P) = -\frac{\Delta H_{vap}}{RT} + C$ $\ln\left(\frac{P_1}{P_2}\right) = \frac{\Delta H_{vap}}{R} \left(\frac{T_1 - T_2}{T_1 T_2}\right)$ $P$ is vapor pressure, $C$ is a constant
Molar heat of sublimation	$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$
Energy required/released during a phase change	$E_{fus} = n\Delta H_{fus}$ $E_{vap} = n\Delta H_{vap}$ $E_{sub} = n\Delta H_{sub}$
Percent by mass (mass fraction)	$w_i = \frac{m(\text{solute})}{m(\text{solution})} \cdot 100\%$
Molality	$b = \frac{n(\text{solute})}{m(\text{solvent})}$ $b$ is molality in mol/kg, $n$ is number of moles, $m$ is mass in kg
Solubility of a gas (Henry's law)	$c = kP$ $k$ is a constant
Vapor pressure of a solution	$P_1 = X_1 P_1^\circ$ $X_1$ is the mole fraction of the <i>solvent</i> in the solution, $P_1^\circ$ is the vapor pressure of the pure solvent
Vapor pressure lowering	$\Delta P = X_2 P_1^\circ$ $X_2$ is the mole fraction of the <i>solute</i> in the solution, $P_1^\circ$ is the vapor pressure of the pure solvent
Boiling point elevation	$\Delta T_b = iK_b b$ $i$ is the van't Hoff factor (see below), $K_b$ is the molal boiling point elevation constant, $b$ is molality

Freezing point depression	$\Delta T_f = iK_f b$ <i>i</i> is the van't Hoff factor (see below), <i>K<sub>f</sub></i> is the molal freezing point depression constant, <i>b</i> is molality
Osmotic pressure	$\pi = iMRT$ <i>i</i> is the van't Hoff factor (see below), <i>M</i> is molarity, <i>R</i> is the gas constant (0.082057 $\frac{\text{L atm}}{\text{K mol}}$ ), <i>T</i> is temperature
van't Hoff factor	$i = \frac{\text{actual number of particles in solution after dissociation}}{\text{number of formula units initially dissolved in solution}}$ In other words, the number of ions each formula unit splits into. This is 1 for all nonelectrolytes. For electrolytes, the real value is always slightly lower than expected because dissociation is never complete.

## Kinetics and equilibrium (Chapter 14 & 15)

Rate law	$\text{rate} = k[A]^x[B]^y$ <i>k</i> is the rate constant, [A] and [B] are reactant concentrations, <i>x</i> and <i>y</i> are the reaction orders of the reactants
Concentration-time for a first-order reaction	$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt$ [A] <sub>t</sub> is the concentration of A at time <i>t</i> , [A] <sub>0</sub> is the concentration of A at time 0
Graphical determination of <i>k</i> for a first-order reaction	$\ln[A]_t = -kt + \ln[A]_0$
Half-life for a first-order reaction	$t_{\frac{1}{2}} = \frac{\ln(2)}{k}$
Concentration-time for a second-order reaction	$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$
Half-life for a second-order reaction	$t_{\frac{1}{2}} = \frac{1}{k[A]_0}$
Concentration-time for a zeroth-order reaction	$[A]_t = -kt + [A]_0$
Half-life for a zeroth-order reaction	$t_{\frac{1}{2}} = \frac{[A]_0}{2k}$

Rate constant from activation energy and temperature (Arrhenius equation)	$k = Ae^{-\frac{E_a}{RT}}$ <p>A is the frequency factor or pre-exponential factor, which can be treated as constant over a wide temperature range. <math>E_a</math> is activation energy, R is the gas constant (<math>8.314 \frac{\text{J}}{\text{K mol}}</math>)</p>
Graphical determination of activation energy	$\ln(k) = \left(-\frac{E_a}{R}\right)\left(\frac{1}{T}\right) + \ln(A)$
Relationship of rate constants at two different temperatures	$\ln\left(\frac{k_1}{k_2}\right) = \frac{E_a}{R}\left(\frac{T_1 - T_2}{T_1 T_2}\right)$
Equilibrium constant (law of mass action)	$K_C = \frac{[C]^c [D]^d \dots}{[A]^a [B]^b \dots}$ <p>C and D are products, and A and B are reactants. Lowercase letters are reaction coefficients. In heterogeneous equilibria, the concentrations of pure solids and liquids are omitted. In homogeneous equilibria, the concentrations of solvents are omitted.</p>
Pressure equilibrium constant	$K_P = \frac{P(C)^c P(D)^d \dots}{P(A)^a P(B)^b \dots}$ <p>P is partial pressure. C and D are gaseous products, and A and B are gaseous reactants. Lowercase letters are reaction coefficients.</p>
Relationship between $K_P$ and $K_C$	$K_P = K_C (RT)^{\Delta n}$ <p>R is the gas constant (<math>0.082057 \frac{\text{L atm}}{\text{K mol}}</math>), <math>\Delta n =</math> moles of gaseous products – moles of gaseous reactants</p>

## Acids and bases (Chapter 16 & 17)

Ion-product constant of water	$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \cdot 10^{-14}$
pH	$\text{pH} = -\log(\text{H}^+)$ $[\text{H}^+] = 10^{-\text{pH}}$
pOH	$\text{pOH} = -\log(\text{OH}^-)$ $[\text{OH}^-] = 10^{-\text{pOH}}$
Relationship between pH and pOH	$\text{pH} + \text{pOH} = 14.00$
Percent ionization	$\text{percent ionization} = \frac{\text{ionized acid concentration at equilibrium}}{\text{initial concentration of acid}} \cdot 100\%$



Relationship between ionization constants of a conjugate acid-base pair	$K_a K_b = K_w$
$pK_a$	$pK_a = -\log(K_a)$ $K_a = 10^{-pK_a}$
pH from $pK_a$ for a buffer solution (Henderson-Hasselbalch equation)	$pH = pK_a + \log\left(\frac{[\text{conjugate base}]}{[\text{acid}]}\right)$
Solubility equilibrium constant / solubility product	$K_{sp} = [A^+]^a [B^-]^b$ where A and B are the ions from the solid being dissolved in the solution, and <i>a</i> and <i>b</i> are their stoichiometric coefficients in the equilibrium equation

## Redox reactions (Chapter 19)

Standard electromotive force (EMF) of a cell	$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$
Free energy change from EMF	$\Delta G = -nFE_{\text{cell}}$ $\Delta G^\circ = -nFE^\circ_{\text{cell}}$ <i>n</i> is the number of moles of electrons transferred in the redox equation <i>F</i> is the Faraday constant, $9.65 \cdot 10^4 \text{ C/mol(e}^-)$
Standard EMF from equilibrium constant	$E^\circ_{\text{cell}} = \frac{RT}{nF} \ln(K)$ At 25 °C/298 K, this takes the form: $E^\circ_{\text{cell}} = \frac{0.0257 \text{ V}}{n} \ln(K)$
EMF under nonstandard conditions (Nernst equation)	$E = E^\circ - \frac{RT}{nF} \ln(Q)$ At 25 °C/298 K, this takes the form: $E = E^\circ - \frac{0.0257 \text{ V}}{n} \ln(Q)$

## Coordination compounds (Chapter 20)

Crystal-field splitting	$\Delta = hf$ where $h$ is Planck's constant ( $6.63 \cdot 10^{-34}$ J·s) and $f$ is the frequency of a photon that causes a d-to-d transition
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