

1 mol = 22.4 L, speed of light =  $3 \times 10^8$  m/s, Planck's constant =  $6.626 \times 10^{-34}$  m<sup>2</sup>kg/s, Avogadro's Constant =  $6.022 \times 10^{23}$  mol<sup>-1</sup>, Faraday = 96,485 C/mol, TGMKhdhdbdcmjnpfa,

R (universal gas constant) =  $8.31446261815324 \frac{J}{Kmol} = 0.082057338 \frac{L \cdot atm}{Kmol}$ ,

Acid/Base equilibrium constants:  $K_a = \frac{[H^+][A^-]}{[HA]}$ ,  $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$ .  $aA + bB \rightarrow cC + dD$ ,  $K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

pH =  $-\log[H^+]$ , pOH =  $-\log[OH^-]$ . pH + pOH = 14 The larger the  $K_a$ , the equilibrium is further to the right, more ions create.  $K_w = [H^+][OH^-] = 10^{-14}$  at 25 degrees C =  $K_a \cdot K_b$ ,  $K_s = [A^+]^a [B^-]^b$

$\Delta G = -nFE^\circ$ , nernst equation:  $E^\circ = \frac{0.0592}{n} \log K$ , Spectroscopy/emission:  $E = h\nu = \frac{hc}{\lambda}$  \*E is in Joules\*, relative

abundance:  $M_1 + M_2(1 - x) = ME$ ,  $Q = Sm\Delta T$ ,  $\Delta G = \Delta H - T\Delta S$ ,  $\Delta S > 0$ : more disorder,  $\Delta G > 0$ : not spontaneous

Beer's Law: absorbance = molar absorptivity \* length of light path \* concentration, property = solute concentration x

constant: boiling point elevation:  $K_b = \frac{R^*T_b^2 M}{\Delta H_{vap}}$ ,  $\Delta T_b = T_b(solution) - T_b(solvent) = K_b * molality$ , freezing point depression

$\Delta T_f = i \cdot K_f \cdot molality$   $i = 1 + a(n - 1)$  n dissociated ions, n=1 i>1. i of fully dissociated ion (strong) = 1.  $K_f$  for water is 1.86 kJmol<sup>-1</sup>

Clausius-Clapeyron  $\ln\left(\frac{P_1}{P_2}\right) = \frac{-\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right)$

## Aqueous Solutions

**Molarity:** mol/L

**Molality:** mol of solute/kg

**Mole Fraction:** Moles of solute/total moles of solution

**Solutes dissolve in solvents to form solutions.**

A Lewis acid: accepts e<sup>-</sup> pair. A Lewis base donates e<sup>-</sup> pair.

A Bronsted Lowry H<sup>+</sup> donor. The conjugate base is formed after the proton is donated. A Bronsted Lowry H<sup>+</sup> acceptor.

The stronger the acid, the weaker the conjugate base. Conjugate bases of strong acids are **spectator ions**.

An Arrhenius acid ↑[H<sup>+</sup>]. An Arrhenius base ↑[OH<sup>-</sup>].

Salts containing Group I elements (Li<sup>+</sup>, Na<sup>+</sup>, K<sup>+</sup>, Cs<sup>+</sup>, Rb<sup>+</sup>) are soluble. There are few exceptions to this rule. Salts

containing the ammonium ion (NH<sub>4</sub><sup>+</sup>) or nitrate ion (NO<sub>3</sub><sup>-</sup>) are also soluble.

Salts containing Cl<sup>-</sup>, Br<sup>-</sup>, or I<sup>-</sup> are generally soluble. Important exceptions to this rule are halide salts of Ag<sup>+</sup>, Pb<sup>2+</sup>, and (Hg<sub>2</sub>)<sup>2+</sup>. Thus, AgCl, PbBr<sub>2</sub>, and Hg<sub>2</sub>Cl<sub>2</sub> are insoluble.

Most silver salts are insoluble. AgNO<sub>3</sub> and Ag(C<sub>2</sub>H<sub>3</sub>O<sub>2</sub>) are common soluble salts of silver; virtually all others are insoluble.

Most sulfate salts are soluble. Important exceptions to this rule include CaSO<sub>4</sub>, BaSO<sub>4</sub>, PbSO<sub>4</sub>, Ag<sub>2</sub>SO<sub>4</sub> and SrSO<sub>4</sub>.

Most hydroxide salts are only slightly soluble. Hydroxide salts of Group I elements are soluble. Calcium and Barium are moderately soluble.

Chromates, phosphates, fluorides, and carbonates are insoluble. Sulfides of transition metals (FeS) are insoluble.

**Colligative properties** are the physical changes that result from adding solute to a solvent. They must accurately describe ideal solutions where enthalpy change is 0 and all intermolecular actions are approximately equal. Colligative properties: boiling-point elevation, freezing-point depression, osmotic pressure (tendency of a solution to take in pure solvent by osmosis), solid-solute solubility, gas-solute solubility.

**Impure substances** tend to have a slightly lower melting point than a pure substance and a broader melting temperature range.

**Suspension** is a heterogeneous mixture of a fluid that contains solid particles large enough for sedimentation.

**Molarity to ppm:** Take molarity and multiply by molar mass to get g/L. Multiply that by 1000 to convert g to mg. PPM is mg/L. Water is polar.

**Unsaturated:** are solutions where the K<sub>sp</sub> has not yet been reached. **Saturated:** K<sub>sp</sub> has been reached.

**Super Saturated** K<sub>sp</sub> has been reached and gone over. These solutions can be achieved by heating up a solvent (heat causes the K<sub>sp</sub> to increase), adding a solute, and then letting the solution cool.

**Miscible** Two things are considered miscible when they can be mixed uniformly in any quantities.

**Raoult's Law:** Vapor pressure of a solution = mole fraction of the solvent\*vapor pressure of the pure solvent. A positive deviation means that there is a higher than expected vapor pressure and is endothermic. A negative deviation means that there is a lower than expected vapor pressure. Ideal solutions follow Raoult's Law.

**Volatile:** Easily evaporated at normal temperatures

**Titration:** a procedure which is used to determine the concentration of an acid or base using volumes or reactants. A measured volume of an acid or base of known concentration is reacted with a sample to the equivalence point.

**Case 1:** Strong base titrated with strong acid. Check mole ratio and set moles of known equal to each other.

**Case 2:** Weak acid titrated with a strong base. At the equivalence point, the reactants of [H+] and [OH-] are gone. There is a slightly basic solution. Calculate pH using the Kb expression.

**Case 3:** Weak base titrated with strong acid. Before equivalence point use Kb, after equivalence point use Ka.

Graph steep vertically at the equivalence point downwards.

Alkaline solutions often feel slippery because they interact with skin oils, resulting in saponification (conversion to soap and alcohol)

**Equivalence point:** Moles Acid = moles Base (using the mole ratio)

**End point:** When indicator turns color. Indicators are chosen so that the endpoint is as close to the equivalence point as possible. **Acid-Base Indicators:** pH Indicators that change color with changing pH. Usually weak acids or bases. KIn is an indicator dissociation constant. Ex: Universal indicator is a mix of indicators that covers a wide pH range. **Thymol Blue - 1st change** - acid color: red, base color: yellow, pH range: 1.2 - 2.8, **Methyl Orange** - acid color: red, base color: yellow, pH range: 3.2 - 4.4, **Bromocresol Green** - acid color: yellow, base color: blue, pH range: 3.8 - 5.4, **Methyl Red** - acid color: yellow, base color: red, pH range: 4.8 - 6.0, **Bromothymol Blue** - acid color: yellow, base color: blue, pH range: 6.0 - 7.6, **Phenol Red** - acid color: yellow, base color: red, pH range: 6.8 - 8.4, **Thymol Blue - 2nd change** - acid color: yellow, base color: blue, pH range: 8.0 - 9.6, **Phenolphthalein** - acid color: colorless, base color: pink, pH range: 8.2 - 10

## Redox Reactions

1. Split it into 2 half-reactions. 2. Balance all non-hydrogen or oxygen elements. Do NOT try to balance one equation with the other, this step ONLY comes at the end. 3. If it's in an acidic medium: Balance oxygen by adding H2O to the appropriate side. Balance hydrogen by adding H+ to the appropriate side. Balance the charge by adding electrons. 4. If it's a basic medium Balance oxygen as above. Balance hydrogen. OH-. add H2O if O unbalanced. Balance the charge as above. 5. Combine: Multiply each reaction by an integer so that there are the same number of electrons on each side

Metal	Oxidation Reaction
Lithium	$\text{Li(s)} \rightarrow \text{Li}^+(\text{aq}) + \text{e}^-$
Potassium	$\text{K(s)} \rightarrow \text{K}^+(\text{aq}) + \text{e}^-$
Barium	$\text{Ba(s)} \rightarrow \text{Ba}^{2+}(\text{aq}) + 2\text{e}^-$
Calcium	$\text{Ca(s)} \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{e}^-$
Sodium	$\text{Na(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{e}^-$
Magnesium	$\text{Mg(s)} \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^-$
Aluminum	$\text{Al(s)} \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{e}^-$
Manganese	$\text{Mn(s)} \rightarrow \text{Mn}^{2+}(\text{aq}) + 2\text{e}^-$
Zinc	$\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$
Chromium	$\text{Cr(s)} \rightarrow \text{Cr}^{3+}(\text{aq}) + 3\text{e}^-$
Iron	$\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$
Cobalt	$\text{Co(s)} \rightarrow \text{Co}^{2+}(\text{aq}) + 2\text{e}^-$
Nickel	$\text{Ni(s)} \rightarrow \text{Ni}^{2+}(\text{aq}) + 2\text{e}^-$
Tin	$\text{Sn(s)} \rightarrow \text{Sn}^{2+}(\text{aq}) + 2\text{e}^-$
Lead	$\text{Pb(s)} \rightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{e}^-$
Hydrogen	$\text{H}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{e}^-$
Copper	$\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$
Silver	$\text{Ag(s)} \rightarrow \text{Ag}^+(\text{aq}) + \text{e}^-$
Mercury	$\text{Hg(l)} \rightarrow \text{Hg}^{2+}(\text{aq}) + 2\text{e}^-$
Platinum	$\text{Pt(s)} \rightarrow \text{Pt}^{2+}(\text{aq}) + 2\text{e}^-$
Gold	$\text{Au(s)} \rightarrow \text{Au}^{3+}(\text{aq}) + 3\text{e}^-$

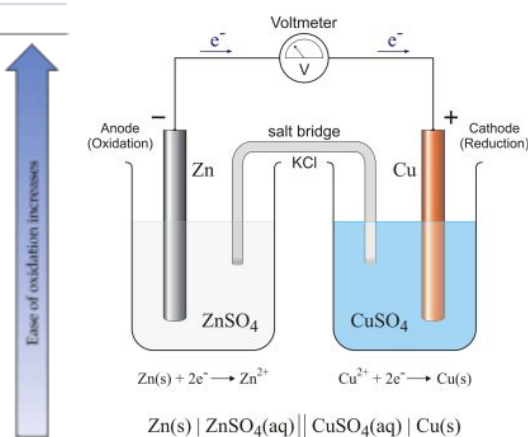


TABLE 17.1 Standard Reduction Potentials at 25 °C		
Reduction Half-Reaction	$E^\circ$ (V)	
$\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq})$	2.87	<div>↑ Weaker oxidizing agent</div> <div>↓ Stronger reducing agent</div>
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	1.78	
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	1.51	
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	1.36	
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	1.33	
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	1.23	
$\text{Br}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$	1.09	
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag(s)}$	0.90	
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	0.77	
$\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2\text{O}_2(\text{aq})$	0.70	
$\text{I}_2(\text{s}) + 2\text{e}^- \rightarrow 2\text{I}^-(\text{aq})$	0.54	
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$	0.40	
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$	0.34	
$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}^{2+}(\text{aq})$	0.15	
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0	
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb(s)}$	-0.13	
$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni(s)}$	-0.26	
$\text{Cd}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cd(s)}$	-0.40	
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe(s)}$	-0.45	
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)}$	-0.76	
$2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$	-0.83	
$\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al(s)}$	-1.66	<div>↑ Stronger oxidizing agent</div> <div>↓ Weaker reducing agent</div>
$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg(s)}$	-2.37	
$\text{Na}^+(\text{aq}) + \text{e}^- \rightarrow \text{Na(s)}$	-2.71	
$\text{Li}^+(\text{aq}) + \text{e}^- \rightarrow \text{Li(s)}$	-3.04	

Table 17.1 Chemistry, 5/e  
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