Electrochemistry Reactions

S.E. Van Bramer, March 14, 1995

Values from Bard, A.; Faulkner, L. *Electrochemical Methods*; Wiley: New York, 1980

Constants

F = 96484.6 C/equiv
R = 8.31441 J Mol^{-1} K^{-1}

Nernst Equation

ox + n e^{1-} <-> red

$$E_{cell} = E_o + \frac{RT}{nF} \ln\left(\frac{[ox]}{[red]}\right)$$

aA + bB + ne^- <-> cC + dD

$$E_{cell} = E_o - \frac{RT}{nF} \ln\left(\frac{[C]^c\ [D]^d}{[A]^a\ [B]^b}\right)$$

$E_{cell} = E_{cathode} - E_{anode}$
using reduction potentials

Comments	E° (V)	Half-reaction	Comments	Trends
F_2 is the strongest oxidizing	3.03	$F_2\,(g) + 2H^+ + 2e^- \rightleftharpoons 2HF$	HF is weakest reducing agent	Spontaneous reaction is
agent	1.3583	$Cl_2\,(g) + 2e^- \rightleftharpoons 2Cl^-$		forward
	1.229	$O_2(g) + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	This reaction limits + potential	(Reduction) for
	1.087	$Br_2\,(g) + 2e^- \rightleftharpoons 2Br^-$	in aqueous solutions.	the Redox couple
	0.905	$2Hg^{2+} + 2e^- \rightleftharpoons Hg_2^{2+}$		(half reaction)
	0.7996	$Ag^+ + e^- \rightleftharpoons Ag$		with more
	0.7961	$Hg_2^{2+} + 2e^- \rightleftharpoons 2Hg$		positive E°.
NOTE: Remember that these	0.770	$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$		
reactions are all	0.522	$Cu^+ + e^- \rightleftharpoons Cu$		Applied potential positive
equilibrium reactions.	0.3402	$Cu^{2+} + 2e^- \rightleftharpoons Cu$		of E° pulls e^-'s off
The may go forward	0.2682	$Hg_2Cl_2 + 2e^- \rightleftharpoons 2Hg + 2Cl^-$		and causes
(reduction) or	0.2415	$Hg_2Cl_2 + 2e^- \rightleftharpoons 2Hg + 2Cl^-$ (sat)	SCE reference electrode	oxidation.
backwards	0.2223	$AgCl(s) + e^- \rightleftharpoons Ag + Cl^-$		
(oxidation).	0.197	$AgCl(s) + e^- \rightleftharpoons Ag + Cl^-$ (sat)	AgCl reference electrode	
Depending upon the	0.15	$Sn^{4+} + 2e^- \rightleftharpoons Sn^{2+}$		
other redox pair	0.0000	$2H^+ + 2e^- \rightleftharpoons H_2$	NHE (SHE) reference	
present.	-0.1263	$Pb^{2+} + 2e^- \rightleftharpoons Pb$	electrode	Applied potential negative
	-0.1364	$Sn^{2+} + 2e^- \rightleftharpoons Sn$		of E° pushes e^-'s
	-0.4026	$Cd^{2+} + 2e^- \rightleftharpoons Cd$		on and causes
	-0.41	$Cr^{3+} + e^- \rightleftharpoons Cr^{2+}$		reduction.
	-0.409	$Fe^{2+} + 2e^- \rightleftharpoons Fe$		
	-0.557	$Cr^{2+} + 2e^- \rightleftharpoons Cr$		
	-0.7628	$Zn^{2+} + 2e^- \rightleftharpoons Zn$		
	-0.8277	$2H_2O + 2e^- \rightleftharpoons H_2 + 2OH^-$	This reaction limits - potential	
	-1.029	$Mn^{2+} + 2e^- \rightleftharpoons Mn$	in aqueous solutions	Spontaneous reaction is
	-1.706	$Al^{3+} + 3e^- \rightleftharpoons Al$ (0.1 M NaOH)		reversed
	-2.375	$Mg^{2+} + 2e^- \rightleftharpoons Mg$		(oxidation) for
	-3.045	$Li^+ + e^- \rightleftharpoons Li$		Redox couple
			Li is strongest reducing agent	(half reaction)
Li^+ is weakest oxidizing agent				with less positive
				E°