TABLE 1 *Standard Reduction Potentials*

Half-cell reaction	Standard electrode potential, E^0 (in volts)	Half-cell reaction	Standard electrode potential, E^0 (in volts)
$F_2 + 2e^- \rightleftarrows 2F^-$	+2.87	$Fe^{3+} + 3e^- \rightleftarrows Fe$	−0.04
$MnO_4^- + 8H^+ + 5e^- \rightleftarrows Mn^{2+} + 4H_2O$	+1.49	$Pb^{2+} + 2e^- \rightleftarrows Pb$	−0.13
$Au^{3+} + 3e^- \rightleftarrows Au$	+1.42	$Sn^{2+} + 2e^- \rightleftarrows Sn$	−0.14
$Cl_2 + 2e^- \rightleftarrows 2Cl^-$	+1.36	$Ni^{2+} + 2e^- \rightleftarrows Ni$	−0.23
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftarrows 2Cr^{3+} + 7H_2O$	+1.33	$Co^{2+} + 2e^- \rightleftarrows Co$	−0.28
$MnO_2 + 4H^+ + 2e^- \rightleftarrows Mn^{2+} + 2H_2O$	+1.21	$Cd^{2+} + 2e^- \rightleftarrows Cd$	−0.40
$Br_2 + 2e^- \rightleftarrows 2Br^-$	+1.07	$Fe^{2+} + 2e^- \rightleftarrows Fe$	−0.41
$Hg^{2+} + 2e^- \rightleftarrows Hg$	+0.85	$S + 2e^- \rightleftarrows S^{2-}$	−0.51
$Ag^+ + e^- \rightleftarrows Ag$	+0.80	$Cr^{3+} + 3e^- \rightleftarrows Cr$	−0.74
$Hg_2^{2+} + 2e^- \rightleftarrows 2Hg$	+0.80	$Zn^{2+} + 2e^- \rightleftarrows Zn$	−0.76
$Fe^{3+} + e^- \rightleftarrows Fe^{2+}$	+0.77	$Al^{3+} + 3e^- \rightleftarrows Al$	−1.66
$MnO_4^- + e^- \rightleftarrows MnO_4^{2-}$	+0.56	$Mg^{2+} + 2e^- \rightleftarrows Mg$	−2.37
$I_2 + 2e^- \rightleftarrows 2I^-$	+0.54	$Na^+ + e^- \rightleftarrows Na$	−2.71
$Cu^{2+} + 2e^- \rightleftarrows Cu$	+0.34	$Ca^{2+} + 2e^- \rightleftarrows Ca$	−2.76
$Cu^{2+} + e^- \rightleftarrows Cu^+$	+0.16	$Ba^{2+} + 2e^- \rightleftarrows Ba$	−2.90
$S + 2H^+(aq) + 2e^- \rightleftarrows H_2S(aq)$	+0.14	$K^+ + e^- \rightleftarrows K$	−2.93
$2H^+(aq) + 2e^- \rightleftarrows H_2$	0.00	$Li^+ + e^- \rightleftarrows Li$	−3.04

CHEMISTRY INTERACTIVE TUTOR

Module 10: Electrochemical Cells

The potential difference across the zinc/hydrogen cell is –0.76 V, so zinc is considered to have an electrode potential of –0.76. The negative number indicates that electrons flow from the zinc electrode, where zinc is oxidized, to the hydrogen electrode, where aqueous hydrogen ions are reduced.

A copper half-cell coupled with the standard hydrogen electrode gives a potential difference measurement of +0.34 V. This positive number indicates that $Cu^{2+}(aq)$ ions are more readily reduced than $H^+(aq)$ ions.

Standard electrode potentials can be used to predict if a redox reaction will occur spontaneously. A spontaneous reaction will have a positive value for E^0_{cell}, which is calculated using the following equation.

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

The half-reaction that has the more negative standard reduction potential will be the anode. Oxidation occurs at the anode, so the anode half-cell reaction will be the reverse of the reduction reaction found in **Table 1.** For this reason, the total potential of a cell is calculated by *subtracting* the standard reduction potential for the reaction at the anode (E^0_{anode}) from the standard reduction potential for the reaction at the cathode ($E^0_{cathode}$).