Homework - p. 670 #1-8

Half-Cells and Cell Potentials

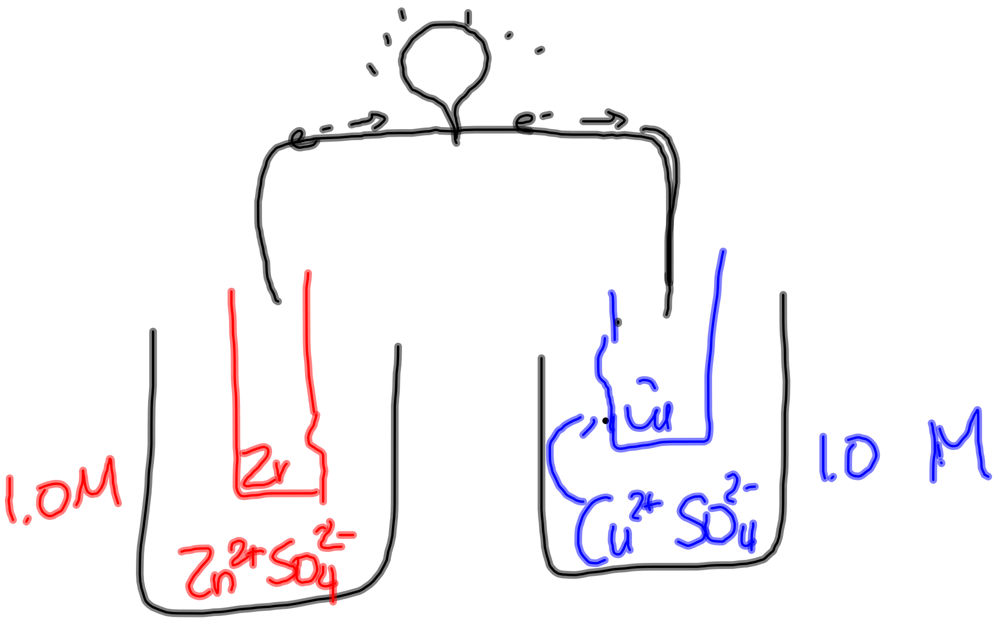
Electrical Potential

- measure of the cell's ability to produce an electric current
- measured in volts (V)
- results from a competition for electrons between two half-cells

reduction potential - tendency of a half-reaction to occur as reduction

cell potential - difference in reduction potentials of the two half-cells

$$\mathbf{E^0_{cell} = E^0_{red} - E^0_{oxid}}$$


e⁻ →
e⁻ →
1.0M
Zn
$Zn^{2+}SO_4^{2-}$
Cu
$Cu^{2+}SO_4^{2-}$
1.0 M

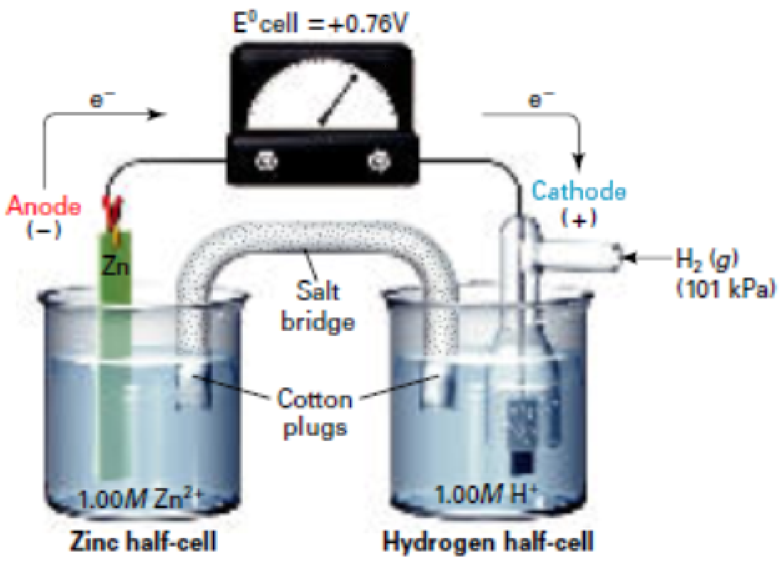
Standard Cell Potential

- measured cell potential in which the ion concentrations in the half-cells at 1M, any gases are at a pressure of 101 kPa, and the temperature is $25^{o}C$.

- a hydrogen electrode serves as a reference point, and it's standard reduction potential is assigned a value of 0.00 V.

Standard Reduction Potentials

- the standard reduction potential for a half-cell can be determined by using a standard hydrogen electrode and the equation for standard cell potential



$$E^0_{cell} = E^0_{red} - E^0_{oxid}$$

Table 21.2

Reduction Potentials at 25°C with 1*M* Concentrations of Aqueous Species

Electrode	Half-reaction	E° (V)
Li^+/Li	$Li^+ + e^- \longrightarrow Li$	−3.05
K^+/K	$K^+ + e^- \longrightarrow K$	−2.93
Ba^{2+}/Ba	$Ba^{2+} + 2e^- \longrightarrow Ba$	−2.90
Ca^{2+}/Ca	$Ca^{2+} + 2e^- \longrightarrow Ca$	−2.87
Na^+/Na	$Na^+ + e^- \longrightarrow Na$	−2.71
Mg^{2+}/Mg	$Mg^{2+} + 2e^- \longrightarrow Mg$	−2.37
Al^{3+}/Al	$Al^{3+} + 3e^- \longrightarrow Al$	−1.66
H_2O/H_2	$2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$	−0.83
Zn^{2+}/Zn	$Zn^{2+} + 2e^- \longrightarrow Zn$	−0.76
Cr^{3+}/Cr	$Cr^{3+} + 3e^- \longrightarrow Cr$	−0.74
Fe^{2+}/Fe	$Fe^{2+} + 2e^- \longrightarrow Fe$	−0.44
H_2O/H_2 (pH 7)	$2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$	−0.42
Cd^{2+}/Cd	$Cd^{2+} + 2e^- \longrightarrow Cd$	−0.40
$PbSO_4/Pb$	$PbSO_4 + 2e^- \longrightarrow Pb + SO_4^{2-}$	−0.36
Co^{2+}/Co	$Co^{2+} + 2e^- \longrightarrow Co$	−0.28
Ni^{2+}/Ni	$Ni^{2+} + 2e^- \longrightarrow Ni$	−0.25
Sn^{2+}/Sn	$Sn^{2+} + 2e^- \longrightarrow Sn$	−0.14
Pb^{2+}/Pb	$Pb^{2+} + 2e^- \longrightarrow Pb$	−0.13
Fe^{3+}/Fe	$Fe^{3+} + 3e^- \longrightarrow Fe$	−0.036
H^+/H_2	$2H^+ + 2e^- \longrightarrow H_2$	0.000
$AgCl/Ag$	$AgCl + e^- \longrightarrow Ag + Cl^-$	+0.22
Hg_2Cl_2/Hg	$Hg_2Cl_2 + 2e^- \longrightarrow 2Hg + 2Cl^-$	+0.27
Cu^{2+}/Cu	$Cu^{2+} + 2e^- \longrightarrow Cu$	+0.34
O_2/OH^-	$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$	+0.40
Cu^+/Cu	$Cu^+ + e^- \longrightarrow Cu$	+0.52
I_2/I^-	$I_2 + 2e^- \longrightarrow 2I^-$	+0.54
Fe^{3+}/Fe^{2+}	$Fe^{3+} + e^- \longrightarrow Fe^{2+}$	+0.77
Hg_2^{2+}/Hg	$Hg_2^{2+} + 2e^- \longrightarrow 2Hg$	+0.79
Ag^+/Ag	$Ag^+ + e^- \longrightarrow Ag$	+0.80
O_2/H_2O (pH 7)	$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+0.82
Hg^{2+}/Hg	$Hg^{2+} + 2e^- \longrightarrow Hg$	+0.85
Br_2/Br^-	$Br_2 + 2e^- \longrightarrow 2Br^-$	+1.07
O_2/H_2O	$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
MnO_2/Mn^{2+}	$MnO_2 + 4H^+ + 2e^- \longrightarrow Mn^{2+} + 2H_2O$	+1.28
$Cr_2O_7^{2-}/Cr^{3+}$	$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$	+1.33
Cl_2/Cl^-	$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
PbO_2/Pb^{2+}	$PbO_2 + 4H^+ + 2e^- \longrightarrow Pb^{2+} + 2H_2O$	+1.46
MnO_4^-/Mn^{2+}	$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$	+1.51
$PbO_2/PbSO_4$	$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \longrightarrow PbSO_4 + 2H_2O$	+1.69
F_2/F^-	$F_2 + 2e^- \longrightarrow 2F^-$	+2.87

Writing the Cell Reaction

$Fe^{3+}_{(aq)} + e^- \Rightarrow Fe^{2+}_{(aq)}$ Red. $E^0_{Fe3+} = 0.77\ V$

$Ni^{2+}_{(aq)} + 2e^- \Leftarrow Ni_{(s)}$ Oxid. $E^0_{Ni2+} = -0.25\ V$

+2 0

Oxidation: $Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + \cancel{2e^-}$

Reduction: ~~$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$~~

$2Fe^{3+}_{(aq)} + \cancel{2e^-} \rightarrow 2Fe^{2+}_{(aq)}$

$Ni_{(s)} + 2Fe^{3+}_{(aq)} \rightarrow Ni^{2+}_{(aq)} + 2Fe^{2+}_{(aq)}$

Calculating the Standard Cell Potential

$Fe^{3+}_{(aq)}$ + e^- $\Rightarrow$ $Fe^{2+}_{(aq)}$ Red. E^0_{Fe3+} = 0.77 V

$Ni^{2+}_{(aq)}$ + $2e^-$ $\Rightarrow$ $Ni_{(s)}$ Ox. E^0_{Ni2+} = -0.25 V

$$E^o_{cell} = E^o_{red} - E^o_{oxid}$$

$$E^o_{cell} = (0.77V) - (-0.25V)$$

$$\boxed{E^o_{cell} = 1.02V}$$

Calculating Standard Cell Potentials

- If the cell potential for a given redox reaction is positive, then the reaction is **spontaneous**.

- If the cell potential is negative then the reaction is **nonspontaneous**.

$$\overset{0}{Zn_{(s)}} + 2\overset{+1}{Ag^{+}_{(aq)}} \quad \overset{+2}{Zn^{2+}_{(aq)}} + 2\overset{0}{Ag_{(s)}}$$

Oxidation: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ $\quad E^\circ = -0.76V$

Reduction: $2Ag^{+}_{(aq)} + 2e^- \rightarrow 2Ag_{(s)}$ $\quad E^\circ = 0.80V$

$$E^\circ_{cell} = E^\circ_{red} - E^\circ_{oxid}$$

$$E^\circ_{cell} = (0.80V) - (-0.76V)$$

$$\boxed{E_{cell} = 1.56V}$$

Spontaneous

$$Zn^{2+}_{(aq)} + 2Ag_{(s)} \longrightarrow Zn_{(s)} + 2Ag^{+}_{(aq)}$$

Oxid.: $2Ag_{(s)} \longrightarrow 2Ag^{+}_{(aq)} + 2e^-$ $\quad$ 0.80V

Red. $Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$ $\quad$ -0.76V

$$E^\circ_{cell} = E^\circ_{red} - E^\circ_{oxid.}$$

$$E^\circ_{cell} = (-0.76V) - (0.80V)$$

$$= -1.52V$$

Homework

p. 675 # 9,10
p. 676 # 11,12
p. 677 # 13-19