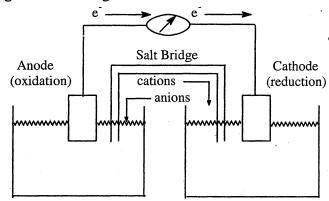
Chapter 11: #17, 18, 19, 26, 29, 32, 39, 46, 53, 55, 56, 91, 95

Galvanic Cells, Cell Potentials, and Standard Reduction Potentials

17. A typical galvanic cell diagram is:



The diagram for all cells will look like this. The contents of each half-cell will be identified for each reaction, with all concentrations at $1.0\,M$ and partial pressures at $1.0\,a$ tm. Note that cations always flow into the cathode compartment and anions always flow into the anode compartment. This is required to keep each compartment electrically neutral.

a. Reference Table 11.1 for standard reduction potentials. Remember that $E_{cell}^{\circ} = E^{\circ}(cathode) - E^{\circ}(anode)$; in the Solutions Guide, we will represent $E^{\circ}(cathode)$ as E_{c}° and represent $-E^{\circ}(anode)$ as $-E_{a}^{\circ}$. Also remember that standard potentials are <u>not</u> multiplied by the integer used to obtain the overall balanced equation.

$$(Cl_2 + 2 e^- \rightarrow 2 Cl^-) \times 3 \qquad \qquad E_c^{\circ} = 1.36 \text{ V}$$

$$7 \text{ H}_2\text{O} + 2 \text{ Cr}^{3+} \rightarrow \text{Cr}_2\text{O}_7^{2-} + 14 \text{ H}^+ + 6 e^- \qquad -E_a^{\circ} = -1.33 \text{ V}$$

$$7 \text{ H}_2\text{O}(1) + 2 \text{ Cr}^{3+}(aq) + 3 \text{ Cl}_2(g) \rightarrow \text{Cr}_2\text{O}_7^{2-}(aq) + 6 \text{ Cl}^-(aq) + 14 \text{ H}^+(aq) \qquad E_{cell}^{\circ} = 0.03 \text{ V}$$

The contents of each compartment is:

Cathode: Pt electrode; Cl₂ bubbled into solution, Cl⁻ in solution

Anode: Pt electrode; Cr³⁺, H⁺, and Cr₂O₇²⁻ in solution

We need a nonreactive metal to use as the electrode in each case, since all of the reactants and products are in solution. Pt is the most common choice. Another possibility is graphite.

b.
$$Cu^{2+} + 2 e^- \rightarrow Cu$$
 $E_c^{\circ} = 0.34 \text{ V}$

$$Mg \rightarrow Mg^{2+} + 2 e^- \qquad -E_a^{\circ} = 2.37 \text{ V}$$

$$Cu^{2+}(aq) + Mg(s) \rightarrow Cu(s) + Mg^{2+}(aq) \qquad E_{cell}^{\circ} = 2.71 \text{ V}$$

Cathode: Cu electrode; Cu²⁺ in solution

Anode: Mg electrode; Mg²⁺ in solution

c.
$$5 e^{-} + 6 H^{+} + IO_{3}^{-} \rightarrow 1/2 I_{2} + 3 H_{2}O$$
 $E_{c}^{\circ} = 1.20 V$

$$(Fe^{2+} \rightarrow Fe^{3+} + e^{-}) \times 5$$
 $-E_{a}^{\circ} = -0.77 V$

$$\frac{6 H^{+} + IO_{3}^{-} + 5 Fe^{2+} \rightarrow 5 Fe^{3+} + 1/2 I_{2} + 3 H_{2}O}{E_{cell}^{\circ} = 0.43 V}$$

or 12 H⁺(aq) + 2 IO₃⁻(aq) + 10 Fe²⁺(aq)
$$\rightarrow$$
 10 Fe³⁺(aq) + I₂(s) + 6 H₂O(1) $E_{cell}^{\circ} = 0.43 \text{ V}$

Cathode: Pt electrode; IO_3^- , I_2 and H_2SO_4 (H⁺ source) in solution.

Anode: Pt electrode; Fe²⁺ and Fe³⁺ in solution

Note: $I_2(s)$ would make a poor electrode since it sublimes.

d.
$$(Ag^{+} + e^{-} \rightarrow Ag) \times 2$$
 $E_{c}^{\circ} = 0.80 \text{ V}$

$$Zn \rightarrow Zn^{2+} + 2 \text{ e} \qquad -E_{a}^{\circ} = 0.76 \text{ V}$$

$$Zn(s) + 2 Ag^{+}(aq) \rightarrow 2 Ag(s) + Zn^{2+}(aq) \qquad E_{cell}^{\circ} = 1.56 \text{ V}$$

Cathode: Ag electrode; Ag+ in solution

Anode: Zn electrode; Zn²⁺ in solution

18. a.
$$2 \text{ Ag}^+ + 2 \text{ e}^- \rightarrow 2 \text{ Ag}$$
 $E_c^\circ = 0.80 \text{ V}$
$$-E_a^\circ = -0.34 \text{ V}$$

$$2 \text{ Ag}^+ (\text{aq}) + \text{Cu(s)} \rightarrow \text{Cu}^{2+} (\text{aq}) + 2 \text{ Ag(s)}$$
 $E_{\text{cell}}^\circ = 0.46 \text{ V}$ Spontaneous at standard conditions $(E_{\text{cell}}^\circ) > 0$.

b.
$$Zn^{2+} + 2 e^- \rightarrow Zn$$
 $E_c^\circ = -0.76 \text{ V}$
 $Ni \rightarrow Ni^{2+} + 2 e^ -E_a^\circ = 0.23 \text{ V}$

$$Zn^{2+}(aq) + Ni(s) \rightarrow Zn(s) + Ni^{2+}(aq) \qquad E_{cell}^\circ = -0.53 \text{ V} \qquad \text{Not spontaneous at standard conditions } (E_{cell}^\circ < 0).$$

c. $(5 e^{-} + 8 H^{+} + MnO_{4}^{-} \rightarrow Mn^{2+} + 4 H_{2}O) \times 2$ $E_{c}^{\circ} = 1.51 V$ $(2 \Gamma \rightarrow I_{2} + 2 e^{-}) \times 5$ $E_{c}^{\circ} = -0.54 V$

$$\frac{16 \text{ H}^{+}(\text{aq}) + 2 \text{ MnO}_{4}^{-}(\text{aq}) + 10 \text{ I}^{-}(\text{aq}) \rightarrow 5 \text{ I}_{2}(\text{s}) + 2 \text{ Mn}^{2+}(\text{aq}) + 8 \text{ H}_{2}\text{O}(1)}{\text{E}_{cell}^{\circ} = 0.97 \text{ V}}$$

Spontaneous since $E_{cell}^{\circ} > 0$.

d.
$$(5 e^{-} + 8 H^{+} + MnO_{4}^{-} \rightarrow Mn^{2+} + 4 H_{2}O) \times 2$$
 $E_{c}^{\circ} = 1.51 V$ $(2 F^{-} \rightarrow F_{2} + 2 e^{-}) \times 5$ $-E_{a}^{\circ} = -2.87 V$

$$16 \text{ H}^{+}(\text{aq}) + 2 \text{ MnO}_{4}^{-}(\text{aq}) + 10 \text{ F}^{-}(\text{aq}) \rightarrow 5 \text{ F}_{2}(\text{g}) + 2 \text{ Mn}^{2+}(\text{aq}) + 8 \text{ H}_{2}\text{O}(\text{l}) \qquad \text{E}_{\text{cell}}^{\circ} = -1.36 \text{ V}$$

Not spontaneous since $E_{cell}^{\circ} < 0$.

19. Reference Exercise 11.17 for a typical galvanic cell design. The contents of each half-cell compartment is identified below with all solute concentrations at 1.0 M and all gases at 1.0 atm. For each pair of half-reactions, the half-reaction with the largest standard reduction potential will be the cathode reaction and the half-reaction with the smallest reduction potential will be reversed to become the anode reaction. Only this combination gives a spontaneous overall reaction, i.e., a reaction with a positive overall standard cell potential.

a.
$$Cl_2 + 2 e^- \rightarrow 2 Cl^ E_c^{\circ} = 1.36 V$$

$$2 Br^{-} \rightarrow Br_2 + 2 e^- \qquad -E_a^{\circ} = -1.09 V$$

$$Cl_2(g) + 2 Br^{\circ}(aq) \rightarrow Br_2(aq) + 2 Cl^{\circ}(aq) \qquad E_{cell}^{\circ} = 0.27 V$$

The contents of each compartment is:

Cathode: Pt electrode; Cl₂(g) bubbled in, Cl⁻ in solution

Anode: Pt electrode; Br₂ and Br in solution

b.
$$(2 e^{-} + 2 H^{+} + IO_{4}^{-} \rightarrow IO_{3}^{-} + H_{2}O) \times 5$$

$$(4 H_{2}O + Mn^{2+} \rightarrow MnO_{4}^{-} + 8 H^{+} + 5 e^{-}) \times 2$$

$$E_{c}^{\circ} = 1.60 \text{ V}$$

$$-E_{a}^{\circ} = -1.51 \text{ V}$$

$$10 H^{+} + 5 IO_{4}^{-} + 8 H_{2}O + 2 Mn^{2+} \rightarrow 5 IO_{3}^{-} + 5 H_{2}O + 2 MnO_{4}^{-} + 16 H^{+}$$

$$E_{cell}^{\circ} = 0.09 \text{ V}$$

This simplifies to:

$$3 H_2O(1) + 5 IO_4^{-1}(aq) + 2 Mn^{2+1}(aq) \rightarrow 5 IO_3^{-1}(aq) + 2 MnO_4^{-1}(aq) + 6 H^{+1}(aq)$$
 $E_{cell}^{\circ} = 0.09 V$

Cathode: Pt electrode; IO₄, IO₃, and H₂SO₄ (as a source of H⁺) in solution

Anode: Pt electrode; Mn²⁺, MnO₄- and H₂SO₄ in solution

c.
$$H_2O_2 + 2 H^+ + 2 e^- \rightarrow 2 H_2O$$
 $E_c^{\circ} = 1.78 V$

$$H_2O_2 \rightarrow O_2 + 2 H^+ + 2 e^- \qquad -E_a^{\circ} = -0.68 V$$

$$\frac{2 H_2O_2(aq) \rightarrow 2 H_2O(l) + O_2(g)}{E_{cell}^{\circ} = 1.10 V}$$

Cathode: Pt electrode; H₂O₂ and H⁺ in solution

Anode: Pt electrode; O₂(g) bubbled in, H₂O₂ and H⁺ in solution

d.
$$(Fe^{3+} + 3 e^{-} \rightarrow Fe) \times 2$$
 $E_c^{\circ} = -0.036 \text{ V}$

$$(Mn \rightarrow Mn^{2+} + 2 e^{-}) \times 3 \qquad -E_a^{\circ} = 1.18 \text{ V}$$

$$\frac{2 Fe^{3+}(aq) + 3 Mn(s) \rightarrow 2 Fe(s) + 3 Mn^{2+}(aq)}{E_{cell}^{\circ} = 1.14 \text{ V}}$$

Cathode: Fe electrode; Fe3+ in solution; Anode: Mn electrode; Mn2+ in solution

26.
$$Cl_2 + 2 e^- \rightarrow 2 Cl$$
 $E^{\circ} = 1.36 \text{ V}$ $Ag^+ + e^- \rightarrow Ag$ $E^{\circ} = 0.80 \text{ V}$ $Pb^{2+} + 2 e^- \rightarrow Pb$ $E^{\circ} = -0.13 \text{ V}$ $Zn^{2+} + 2 e^- \rightarrow Zn$ $E^{\circ} = -0.76 \text{ V}$ $Na^+ + e^- \rightarrow Na$ $E^{\circ} = -2.71 \text{ V}$

- a. Oxidizing agents (species reduced) are on the left side of the above reduction half-reactions. Of the species available, Ag⁺ would be the best oxidizing agent since it has the most positive E° value.
- b. Reducing agents (species oxidized) are on the right side of the reduction half-reactions. Of the species available, Zn would be the best reducing agent since it has the most positive -E° value.
- c. $SO_4^{2-} + 4 H^+ + 2 e^- \rightarrow H_2SO_3 + H_2O$ $E_c^{\circ} = 0.20 \text{ V}$; SO_4^{2-} can oxidize Pb and Zn at standard conditions. When SO_4^{2-} is coupled with these reagents, E_{cell}° is positive.
- d. Al \rightarrow Al³⁺ + 3 e⁻ E_a^o = 1.66 V; Al can oxidize Ag⁺ and Zn²⁺ at standard conditions since E_{cell} > 0.

29.
$$ClO^- + H_2O + 2 e^- \rightarrow 2 OH^- + Cl^ E^\circ = 0.90 \text{ V}$$

 $2 \text{ NH}_3 + 2 OH^- \rightarrow N_2H_4 + 2 H_2O + 2 e^ -E^\circ = 0.10 \text{ V}$
 $ClO^-(aq) + 2 \text{ NH}_3(aq) \rightarrow Cl^-(aq) + N_2H_4(aq) + H_2O(l)$ $E_{cell}^\circ = 1.00 \text{ V}$

Since E_{cell}° is positive for this reaction, then at standard conditions ClO can spontaneously oxidize NH₃ to the somewhat toxic N₂H₄.

32. Consider the strongest oxidizing agent combined with the strongest reducing agent from Table 11.1:

$$F_2 + 2 e^{-} \rightarrow 2 F^{-}$$
 $E_c^{\circ} = 2.87 V$
 $(Li \rightarrow Li^{+} + e^{-}) \times 2$ $-E_a^{\circ} = 3.05 V$
 $F_2(g) + 2 Li(s) \rightarrow 2 Li^{+}(aq) + 2 F^{-}(aq)$ $E_{cell}^{\circ} = 5.92 V$

The claim is impossible. The strongest oxidizing agent and reducing agent when combined only give E_{cell}° of about 6 V.

$$\Delta G^{\circ} = -nFE^{\circ} = \Delta H^{\circ} - T\Delta S^{\circ}, E^{\circ} = \frac{T\Delta S^{\circ}}{nF} - \frac{\Delta H^{\circ}}{nF}$$

If we graph E° vs. T we should get a straight line (y = mx + b). The slope of the line (m) is equal to $\Delta S^{\circ}/nF$ and the y-intercept (b) is equal to $\Delta H^{\circ}/nF$.

46.
$$Ag^{+} + e^{-} \rightarrow Ag \qquad \qquad E_{c}^{\circ} = 0.80 \text{ V}$$

$$Ag + 2 S_{2}O_{3}^{2-} \rightarrow Ag(S_{2}O_{3})_{2}^{3-} + e^{-} \qquad -E_{a}^{\circ} = -0.017 \text{ V}$$

$$\overline{Ag^{+}(aq) + 2 S_{2}O_{3}^{2-}(aq) \rightarrow Ag(S_{2}O_{3})_{2}^{3-}(aq)} \qquad E_{cell}^{\circ} = 0.78 \text{ V} \qquad K = ?$$
For this overall reaction,
$$E_{cell}^{\circ} = \frac{0.0591}{n} \log K$$

$$\log K = \frac{nE^{\circ}}{0.0591} = \frac{(1)(0.78)}{0.0591} = 13.20, \quad K = 10^{13.20} = 1.6 \times 10^{13}$$

Galvanic Cells: Concentration Dependence

53. a. n = 2 for this reaction (lead goes from Pb \rightarrow Pb²⁺ in PbSO₄).

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \left(\frac{1}{[\text{H}^{+}]^{2} [\text{HSO}_{4}^{-}]^{2}} \right) = 2.04 \text{ V} - \frac{0.0591}{2} \log \frac{1}{(4.5)^{2} (4.5)^{2}}$$

$$2.04 \text{ V} - (-0.077 \text{ V}) = 2.12 \text{ V}$$

b. We can calculate ΔG° from $\Delta G^{\circ} = \Delta H^{\circ}$ - $T\Delta S^{\circ}$ and then E° from $\Delta G^{\circ} = -nFE^{\circ}$; or we can use the equation derived in Exercise 11.39.

$$E_{-20}^{\circ} = \frac{T\Delta S^{\circ} - \Delta H^{\circ}}{nF} = \frac{(253 \text{ K}) (263.5 \text{ J/K}) + 315.9 \times 10^{3} \text{ J}}{(2 \text{ mol e}^{-}) (96,485 \text{ C/mol e}^{-})} = 1.98 \text{ J/C} = 1.98 \text{ V}$$

c.
$$E_{-20} = E_{-20}^{\circ} - \frac{RT}{nF} \ln Q = 1.98 \text{ V} - \frac{RT}{nF} \ln \frac{1}{[H^+]^2 [HSO_4^-]^2}$$

$$E_{-20} = 1.98 \text{ V} - \frac{(8.3145 \text{ J K}^{-1} \text{ mol}^{-1}) (253 \text{ K})}{(2 \text{ mol e}^{-}) (96,485 \text{ C/mol e}^{-})} \ln \frac{1}{(4.5)^2 (4.5)^2} = 1.98 \text{ V} - (-0.066 \text{ V}) = 2.05 \text{ V}$$

- d. As the temperature decreases, the cell potential decreases. Also, oil becomes more viscous at lower temperatures, which adds to the difficulty of starting an engine on a cold day. The combination of these two factors results in batteries failing more often on cold days than on warm days.
- As is the case for all concentration cells, $E_{\text{cell}}^{\circ} = 0$, and the smaller ion concentration is always in the anode compartment. The general Nernst equation for the Ni | Ni²⁺ (x M) | Ni²⁺ (y M) | Ni concentration cell is:

$$E_{cell} = E_{cell}^{\circ} - \frac{0.0591}{n} \log Q = \frac{-0.0591}{2} \log \frac{\left[Ni^{2+}\right]_{anode}}{\left[Ni^{2+}\right]_{cathode}}$$

- a. Since both compartments are at standard conditions ([Ni²⁺] = 1.0 M), then $E_{cell} = E_{cell}^{\circ} = 0 \text{ V}$. No electron flow occurs.
- b. Cathode = $2.0 M \text{ Ni}^{2+}$; Anode = $1.0 M \text{ Ni}^{2+}$; Electron flow is always from the anode to the cathode, so electrons flow to the right in the diagram.

$$E_{\text{cell}} = \frac{-0.0591}{2} \log \frac{[\text{Ni}^{2+}]_{\text{anode}}}{[\text{Ni}^{2+}]_{\text{cethode}}} = \frac{-0.0591}{2} \log \frac{1.0}{2.0} = 8.9 \times 10^{-3} \text{ V}$$

c. Cathode = $1.0 M \text{ Ni}^{2+}$; Anode = $0.10 M \text{ Ni}^{2+}$; Electrons flow to the left in the diagram.

$$E_{cell} = \frac{-0.0591}{2} \log \frac{0.10}{1.0} = 0.030 \text{ V}$$

d. Cathode = $1.0 M \text{ Ni}^{2+}$; Anode = $4.0 \times 10^{-5} M \text{ Ni}^{2+}$; Electrons flow to the left in the diagram.

$$E_{cell} = \frac{-0.0591}{2} \log \frac{4.0 \times 10^{-5}}{1.0} = 0.13 \text{ V}$$

e. Since both concentrations are equal, $\log (2.5/2.5) = \log 1.0 = 0$ and $E_{cell} = 0$. No electron flow occurs.

$$M^{2+} + 2e^- \rightarrow M(s)$$

$$(a) \rightarrow M^{2+} + 2a^{-}$$

$$E_{c}^{o} = 0.80 \text{ V}$$

$$M^{2+}$$
 (cathode) $\rightarrow M^{2+}$ (anode)

$$E_{cell}^{\circ} = 0.00 \text{ V}$$

node:
$$M^{2+} + 2e^{-} \rightarrow M(s)$$
 $E_{c}^{\circ} = 0.80 \text{ V}$
de: $M(s) \rightarrow M^{2+} + 2e^{-}$ $-E_{a}^{\circ} = -0.80 \text{ V}$

$$M^{2+} \text{ (cathode)} \rightarrow M^{2+} \text{ (anode)}$$
 $E_{cell}^{\circ} = 0.00 \text{ V}$

$$E_{cell} = 0.44 \text{ V} = 0.00 \text{ V} - \frac{0.0591}{2} \log \frac{[M^{2+}]_{anode}}{[M^{2+}]_{cathode}}, 0.44 = -\frac{0.0591}{2} \log \frac{[M^{2+}]_{anode}}{1.0}$$

$$\log [M^{2+}]_{\text{anode}} = -\frac{2(0.44)}{0.0591} = -14.89, [M^{2+}]_{\text{anode}} = 1.3 \times 10^{-15} M$$

Since we started with equal numbers of moles of SO_4^{2-} and M^{2+} , then $[M^{2+}] = [SO_4^{2-}]$ at equilibrium.

$$K_{sp} = [M^{2+}][SO_4^{2-}] = (1.3 \times 10^{-15})^2 = 1.7 \times 10^{-30}$$

Challenge Problems

91. a.
$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$
 $E_{cell}^{\circ} = 1.10 \text{ V}; E_{cell} = 1.10 \text{ V} - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$

$$E_{cell} = 1.10 \text{ V} - \frac{0.0591}{2} \log \frac{0.10}{2.50} = 1.10 \text{ V} - (-0.041 \text{ V}) = 1.14 \text{ V}$$

b.
$$10.0 \text{ h} \times \frac{60 \text{ min}}{\text{h}} \times \frac{60 \text{ s}}{\text{min}} \times \frac{10.0 \text{ C}}{\text{s}} \times \frac{1 \text{ mol e}}{96,485 \text{ C}} \times \frac{1 \text{ mol Cu}}{2 \text{ mol e}} = 1.87 \text{ mol Cu produced}$$

The Cu²⁺ concentration decreases by 1.87 mol/L and the Zn²⁺ concentration will increase by 1.87 mol/L.

$$[Cu^{2+}] = 2.50 - 1.87 = 0.63 M; [Zn^{2+}] = 0.10 + 1.87 = 1.97 M$$

$$E_{cell} = 1.10 \text{ V} - \frac{0.0591}{2} \log \frac{1.97}{0.63} = 1.10 \text{ V} - 0.015 \text{ V} = 1.09 \text{ V}$$

c. 1.87 mol Zn consumed
$$\times \frac{65.38 \text{ g Zn}}{\text{mol Zn}} = 122 \text{ g Zn}$$
; Mass of electrode = 200. - 122 = 78 g Zn

1.87 mol Cu formed
$$\times \frac{63.55 \text{ g Cu}}{\text{mol Cu}} = 119 \text{ g Cu}$$
; Mass of electrode = 200. + 119 = 319 g Cu

- d. Three things could possibly cause this battery to go dead:
 - 1. All of the Zn is consumed.
 - 2. All of the Cu²⁺ is consumed.
 - 3. Equilibrium is reached ($E_{cell} = 0$).

We began with 2.50 mol Cu^{2+} and 200. g $Zn \times 1$ mol Zn/65.38 g Zn = 3.06 mol Zn. Cu^{2+} is the limiting reagent and will run out first. To react all the Cu²⁺ requires:

2.50 mol Cu²⁺ ×
$$\frac{2 \text{ mol e}}{\text{mol Cu}^{2+}}$$
 × $\frac{96,485 \text{ C}}{\text{mol e}}$ × $\frac{1 \text{ s}}{10.0 \text{ C}}$ × $\frac{1 \text{ h}}{3600 \text{ s}}$ = 13.4 h

For equilibrium to be reached: $E = 0 = 1.10 \text{ V} - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Cn^{2+}]}$

$$\frac{[Zn^{2+}]}{[Cu^{2+}]} = K = 10^{2(1.10)/0.0591} = 1.68 \times 10^{37}$$

This is such a large equilibrium constant that virtually all of the Cu²⁺ must react to reach equilibrium. So, the battery will go dead in 13.4 hours.

$$3 \times (e^{-} + 2 H^{+} + NO_{3}^{-} \rightarrow NO_{2} + H_{2}O)$$
 $E_{c}^{\circ} = 0.775 V$
 $2 H_{2}O + NO \rightarrow NO_{3}^{-} + 4 H^{+} + 3 e^{-}$ $-E_{a}^{\circ} = -0.957 V$

$$E_c = 0.775 \text{ V}$$

 $E_a^\circ = -0.957 \text{ V}$

$$\overline{2 \text{ H}^{+}(\text{aq}) + 2 \text{ NO}_{3}^{-}(\text{aq}) + \text{NO}(g)} \rightarrow 3 \text{ NO}_{2}(g) + \text{H}_{2}\text{O}(1)$$
 $\text{E}_{\text{cell}}^{\circ} = -0.182 \text{ V}$ $\text{K} = ?$

$$E_{cell}^{\circ} = -0.182 \text{ V } K = ?$$

$$\log K = \frac{nE^{\circ}}{0.0591} = \frac{3(-0.182)}{0.0591} = -9.239, K = 10^{-9.239} = 5.77 \times 10^{-10}$$

b. Let
$$C = \text{concentration of } HNO_3 = [H^+] = [NO_3^-]$$

$$5.77 \times 10^{-10} = \frac{P_{\text{NO}_2}^3}{P_{\text{NO}} \times [\text{H}^+]^2 \times [\text{NO}_3^-]^2} = \frac{P_{\text{NO}_2}^3}{P_{\text{NO}} \times C^4}$$

If 0.20 mol % NO_2 and $P_{tot} = 1.00$ atm:

$$P_{NO_2} = \frac{0.20 \text{ mol NO}_2}{100. \text{ mol total}} \times 1.00 \text{ atm} = 2.0 \times 10^{-3} \text{ atm}; \ P_{NO} = 1.00 - 0.0020 = 1.00 \text{ atm}$$

$$5.77 \times 10^{-10} = \frac{(2.0 \times 10^{-3})^3}{(1.00) \text{ C}^4}$$
, C = 1.9 M HNO₃