17.18 (a) $2[Al \longrightarrow Al^{3+} + 3e^-]$ $\quad E^\circ_{ox} = -1.185\,V$

$3[2e^- + Mn^{2+} \longrightarrow Mn]$ $\quad E^\circ_{red} = -1.185\,V$

$2Al + 3Mn^{2+} \longrightarrow 2Al^{3+} + 3Mn$ $\quad E^\circ = 0.477\,V$

$\Delta G^\circ = -nFE^\circ = -6\ mole\ e^- \times \frac{96485\,C}{mole\ e^-} \times \frac{0.477\,J}{C} = -2.76 \times 10^5\,J$ spontaneous

$K = \exp\left(-\frac{\Delta G^\circ}{RT}\right) = e^{\exp\left(\frac{2.76 \times 10^5\,J}{8.3145 \times 298}\right)} = 2 \times 10^{48}$ effectively complete reaction

(b) $2Hg \longrightarrow Hg_2^{2+} + 2e^-$ $\quad E^\circ_{ox} = -.7973\,V$

$2e^- + Zn^{2+} \longrightarrow Zn$ $\quad E^\circ_{red} = -0.7618\,V$

$2Hg + Zn^{2+} \longrightarrow Hg_2^{2+} + Zn$ $\quad E^\circ = -1.559\,V$

$\Delta G^\circ = -nFE^\circ = -2\ mole\ e^- \times \frac{96485\,C}{mole\ e^-} \times -\frac{1.559\,J}{C} = 3.0 \times 10^5\,J$ not spontaneous

$K = 2 \times 10^{-53}$ effectively 0

17.19 For a spontaneous reaction $E^\circ > 0$. Silver is oxidized. Gold is reduced.

(a) $Au^{3+} + 3e^- \longrightarrow Au$ $\quad E^\circ_{red} = 1.40\,V$ CATHODE

$3Ag \longrightarrow 3Ag^+ + 3e^-$ $\quad E^\circ_{ox} = -E^\circ_{red} = -0.80\,V$ ANODE

$E^\circ = 0.60\,V$

(b) $Au^{3+} + 3Ag \longrightarrow Au + 3Ag^+$

(c) $E^\circ = 0.60\,V$ $\quad \Delta G^\circ = -nFE^\circ = 3\ mole\ e^- \times \frac{96485\,C}{mole\ e^-} \times \frac{0.60\,J}{C} = -1.7 \times 10^2\,kJ$ spontaneous

$K = 3 \times 10^{30}$ so effectively complete reaction