# Thermodynamics of electrodes and Galvanic cells

P57. Calculate the mean activity coefficient of a 0.50 mol dm$^{-3}$ aqueous lead(II) nitrate solution from the fact that the electrode potential of a lead electrode immersed into this solution is –0,158 V at 25.0 °C and $E^{\ominus}(Pb^{2+}/Pb)$ = –0.130 V. *[$\gamma_{\pm}$ = 0.226]*

P58. The cell potential of the following galvanic cell is 0.2848 V at 20.0 °C:

$$Ag(s)\ |\ 0.010\ M\ AgNO_3(aq) + 0.10\ M\ NH_3(aq)\ ||\ 0.010\ M\ AgNO_3(aq)\ |\ Ag(s)$$

The mean activity coefficient in the right-hand side solution is $\gamma_{\pm}$ = 0.896. Calculate the net stability constant of the following reaction: $Ag^+ + 2\ NH_3 \rightleftharpoons Ag(NH_3)_2^+$. *[$K = 1.37\times10^7$]*

P59. The cell potential of the following galvanic cell is 0.8356 V at 25.0 °C:

$$Pt(H_2(g, 1\ bar))\ |\ HClO_4\ (aq, a = 1)\ ||\ Hg(ClO_4)_2\ (aq, a = 6{,}0\cdot10^{-3})\ |\ Hg_2(ClO_4)_2\ (aq, a = 6{,}0\cdot10^{-3})\ |\ Pt(s)$$

Calculate the standard electrode potential of the $Hg^{2+}/Hg_2^{2+}$ half reaction. *[$E^{\ominus}$= 0.901 V]*

P60. The cell potential of the following galvanic cell is 0. 9647 V at 25.0 °C:

$$Pb(s)\ |\ PbSO_4(aq)\ |\ Na_2SO_4(aq)\ |\ Hg_2SO_4(aq)\ |\ Hg(s)$$

The temperature coefficient of the cell potential is $1.74\cdot10^{-4}$ V K$^{-1}$. Give the chemical reaction responsible for the production of electricity, and calculate its reaction heat and reaction Gibbs free energy. [$Pb(s) + Hg_2SO_4(aq) \rightarrow PbSO_4(aq) + 2Hg(s)$; *$\Delta_rG$ = –186.2 kJ/mol; $\Delta_rH$ = –176.1 kJ/mol]*

P61. Calculate the equilibrium constant of the following reaction at 25.0 °C:

$$Fe^{3+} + Cu^+ \rightleftharpoons Fe^{2+} + Cu^{2+}$$

$E^{\ominus}(Fe^{3+}/Fe^{2+})$ = 0.77 V and $E^{\ominus}(Cu^{2+}/Cu^+)$ = 0.17 V. *[$K = 1.39\times10^{10}$]*

P62. Calculate the equilibrium constant of the following reaction at 20.0 °C:

$$2\ FeCl_2 + H_3AsO_4 + 2\ HCl \rightleftharpoons 2\ FeCl_3 + H_3AsO_3 + H_2O$$

$E^{\ominus}(Fe^{3+}/Fe^{2+})$ = 0.772 V and $E^{\ominus}(AsO_4^{3-}/AsO_3^{3-})$ = 0.630 V (the electrode half-reaction in the second case is as follows: $AsO_4^{3-} + 2H^+ + 2e^- \longrightarrow AsO_3^{3-} + H_2O$. *[$K = 1.31\times10^{-5}$]*

P63. Using the fact that the Gibbs free energy is a state function, calculate $E^{\ominus}(Cu^+/Cu)$ from the following standard electrode potential values: $E^{\ominus}(Cu^{2+}/Cu)$ = 0.34 V and $E^{\ominus}(Cu^{2+}/Cu^+)$ = 0.16 V. *[$E^{\ominus}(Cu^+/Cu)$ = 0.52 V]*