13-14. What is the difference between E and E° for a redox reaction? Which one runs down to 0 when the complete cell comes to equilibrium?

13-19. Suppose that the concentrations of NaF and KCl were each 0.10 M in the cell

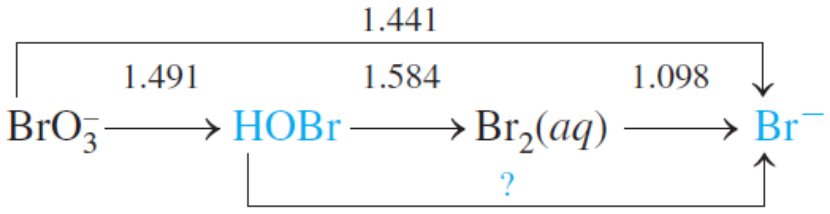
$$\mathrm{Pb}(s)\,|\,\mathrm{PbF_2}(s)\,|\,\mathrm{F^-}(aq)\,\|\,\mathrm{Cl^-}(aq)\,|\,\mathrm{AgCl}(s)\,|\,\mathrm{Ag}(s)$$

(a) Using the half-reactions $2\mathrm{AgCl}(s) + 2e^- \rightleftharpoons 2\mathrm{Ag}(s) + 2\mathrm{Cl^-}$ and $\mathrm{PbF_2}(s) + 2e^- \rightleftharpoons \mathrm{Pb}(s) + 2\mathrm{F^-}$, calculate the cell voltage.

(b) By the reasoning in Figure 13-8, in which direction do electrons flow?

(c) Now calculate the cell voltage by using the reactions $2\mathrm{Ag^+} + 2e^- \rightleftharpoons 2\mathrm{Ag}(s)$ and $\mathrm{Pb^{2+}} + 2e^- \rightleftharpoons \mathrm{Pb}(s)$. For this part, you will need the solubility products for $\mathrm{PbF_2}$ and AgCl.

13-21. Write a balanced chemical equation (in acidic solution) for the reaction represented by the question mark on the lower arrow.[18] Calculate E° for the reaction.



13-26. A solution contains 0.100 M Ce^{3+}, 1.00×10^{-4} M Ce^{4+}, 1.00×10^{-4} M Mn^{2+}, 0.100 M MnO_4^-, and 1.00 M $HClO_4$.

(a) Write a balanced net reaction that can occur between species in this solution.

(b) Calculate ΔG° and K for the reaction.

(c) Calculate E for the conditions given.

(d) Calculate ΔG for the conditions given.

(e) At what pH would the given concentrations of Ce^{4+}, Ce^{3+}, Mn^{2+}, and MnO_4^- be in equilibrium at 298 K?

14-2. Convert the following potentials. The Ag | AgCl and calomel reference electrodes are saturated with KCl.

(a) 0.523 V versus S.H.E. = ? versus Ag | AgCl

(b) −0.111 V versus Ag | AgCl = ? versus S.H.E.

(c) −0.222 V versus S.C.E. = ? versus S.H.E.

(d) 0.023 V versus Ag | AgCl = ? versus S.C.E.

(e) −0.023 V versus S.C.E. = ? versus Ag | AgCl

14-6. A cell was prepared by dipping a Cu wire and a saturated calomel electrode into 0.10 M $CuSO_4$ solution. The Cu wire was attached to the positive terminal of a potentiometer and the calomel electrode was attached to the negative terminal.

(a) Write a half-reaction for the Cu electrode.

(b) Write the Nernst equation for the Cu electrode.

(c) Calculate the cell voltage.

14-8. A 10.0-mL solution of 0.050 0 M $AgNO_3$ was titrated with 0.025 0 M NaBr in the cell

$$\text{S.C.E.} \parallel \text{titration solution} \mid \text{Ag}(s)$$

Find the cell voltage for 0.1 and 30.0 mL of titrant.

14-33. A cyanide ion-selective electrode obeys the equation

$$E = \text{constant} - 0.059\,16 \log[CN^-]$$

The potential was −0.230 V when the electrode was immersed in 1.00 mM NaCN.

(a) Evaluate the constant in the equation.

(b) Using the result from part **(a)**, find $[CN^-]$ if $E = -0.300$ V.

(c) Without using the constant from part **(a)**, find $[CN^-]$ if $E = -0.300$ V.

14-34. By how many volts will the potential of an ideal Mg^{2+} ion-selective electrode change if the electrode is removed from 1.00×10^{-4} M $MgCl_2$ and placed in 1.00×10^{-3} M $MgCl_2$ at 25°C?