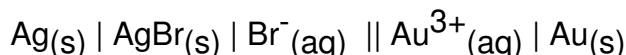


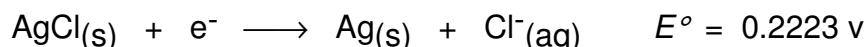
Review Problems – Chapter 19

(Answers on next page)

- (1) A large electrolytic cell that produces metallic aluminum from Al_2O_3 ore is capable of making 250 kg of aluminum in 24 hours. Determine the current (in amps) that is required for this process. Include appropriate chemical reactions.
- (2) An aqueous solution of NaCl was electrolyzed with a current of 2.50 amps for 15.0 minutes. What volume (in mL) of 0.500 M HCl would be required to neutralize the resulting solution? (*Hint:* H_2 is produced at the cathode and Cl_2 at the anode.)
- (3) Under standard conditions, is the following a galvanic or an electrolytic cell? Support your conclusion with appropriate calculations and balanced chemical reactions.

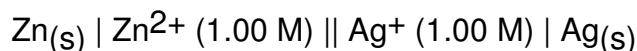


- (4) A silver wire coated with AgCl is sensitive to the chloride ion concentration because of the following half-cell reaction.

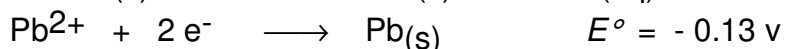
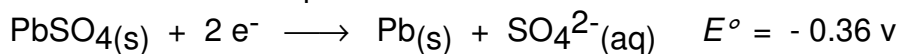


Calculate the molar concentration of Cl^- when the potential of this half-cell is measured to be 0.4900 volts relative to a standard hydrogen electrode.

- (5) Consider the following electrochemical cell in which the volume of solution in each half-cell is 100 mL.

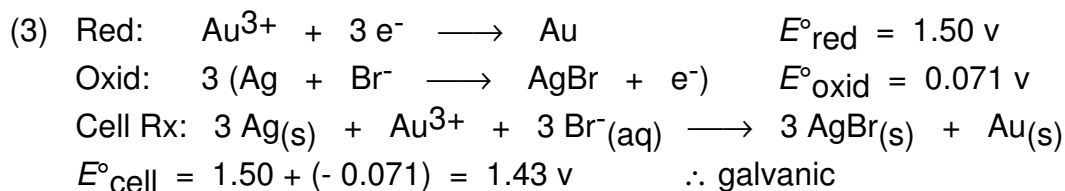


- (a) Write balanced chemical equations for the anode, cathode, and overall cell reactions.
- (b) Determine E_{cell}° , ΔG° , and the equilibrium constant (K_{C}) for the cell reaction.
- (c) If current is drawn from this cell at a constant rate of 0.10 amp, what will the cell potential be after 24.0 hours?
- (6) Given the following standard reduction potentials, calculate the solubility product constant (K_{Sp}) for lead sulfate, PbSO_4 .



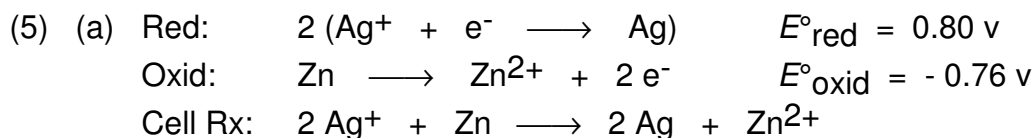
$$\begin{aligned}
 (1) \quad & (250,000 \text{ g}) (1 \text{ mole Al} / 26.95 \text{ g}) (3 \text{ mole } e^- / 1 \text{ mole Al}) (96,500 \text{ coul} / \text{mole } e^-) \\
 & = 2.686 \times 10^9 \text{ coul} = 2.686 \times 10^9 \text{ amp}\cdot\text{sec} \\
 & (2.686 \times 10^9 \text{ amp}\cdot\text{sec}) / (24 \text{ hr}) (3600 \text{ sec/hr}) = 3.11 \times 10^4 \text{ amp}
 \end{aligned}$$

$$\begin{aligned}
 (2) \quad \text{Cathode Rx: } & 2 \text{ H}_2\text{O} + 2 e^- \longrightarrow \text{H}_2 + 2 \text{ OH}^- \\
 & (2.50 \text{ amp}) (15 \text{ min}) (60 \text{ sec/min}) = 2250 \text{ amp}\cdot\text{sec} = 2250 \text{ coul} \\
 & (2250 \text{ coul}) (1 \text{ mole } e^- / 96500 \text{ coul}) (2 \text{ mole OH}^- / 2 \text{ mole } e^-) \\
 & = 0.02332 \text{ mole OH}^- = 0.02332 \text{ mole HCl} \\
 & (0.02332 \text{ mole HCl}) (1000 \text{ mL} / 0.500 \text{ mole HCl}) = 46.6 \text{ mL HCl}
 \end{aligned}$$



(4) Apply the Nernst equation to the given half-cell:

$$\begin{aligned}
 E_{\text{cell}} &= E^\circ_{\text{cell}} - (0.0592 / n) \log Q \\
 0.4900 \text{ v} &= 0.2223 \text{ v} - (0.0592 / 1) \log Q \\
 Q &= [\text{Cl}^-] = 3.0 \times 10^{-5} \text{ M}
 \end{aligned}$$



$$(b) \quad E^\circ_{\text{cell}} = 0.80 + 0.76 = 1.56 \text{ v}$$

$$\begin{aligned}
 \Delta G^\circ &= -nFE^\circ_{\text{cell}} = -(2 \text{ moles}) (96,500 \text{ coul/mole}) (1.56 \text{ J/coul}) \\
 &= -301,000 \text{ J} = -301 \text{ kJ}
 \end{aligned}$$

$$\begin{aligned}
 \log K_{\text{C}} &= nE^\circ_{\text{cell}} / 0.0592 = (2)(1.56) / 0.0592 = 52.7 \\
 K_{\text{C}} &= 5.0 \times 10^{52}
 \end{aligned}$$

$$(c) \quad (0.10 \text{ amp}) (24 \text{ hr}) (3600 \text{ sec/hr}) = 8640 \text{ amp}\cdot\text{sec} = 8640 \text{ coul}$$

$$(8640 \text{ coul}) (1 \text{ mole } e^- / 96,500 \text{ coul}) = 0.0895 \text{ mole } e^- \text{ transferred}$$

$$\text{mole Ag}^+ \text{ consumed} = (0.0895 \text{ mole } e^-) (1 \text{ mole Ag}^+ / 1 \text{ mole } e^-)$$

$$= 0.0895 \text{ mole Ag}^+$$

$$\text{mole Ag}^+ \text{ remaining (after 24 hrs)} = 0.100 - 0.0895 = 0.0105 \text{ mole Ag}^+$$

$$[\text{Ag}^+] = 0.0105 \text{ mole} / 0.100 \text{ L} = 0.105 \text{ M}$$

$$\text{mole Zn}^{2+} \text{ produced} = (0.0895 \text{ mole } e^-) (1 \text{ mole Zn}^{2+} / 2 \text{ mole } e^-)$$

$$= 0.04475 \text{ mole}$$

$$\text{mole Zn}^{2+} \text{ after 24 hrs} = 0.100 + 0.04475 = 0.145 \text{ mole}$$

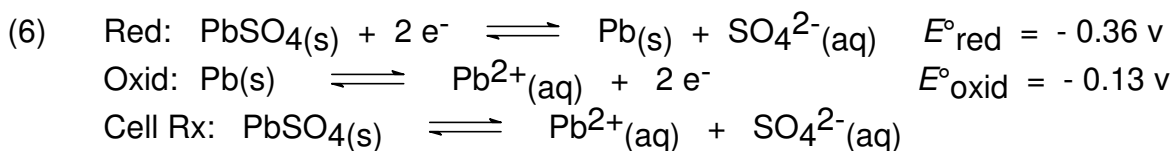
$$[\text{Zn}^{2+}] = 0.145 \text{ mole} / 0.100 \text{ L} = 1.45 \text{ M}$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - (0.0592 / n) \log Q$$

$$Q = [\text{Zn}^{2+}] / [\text{Ag}^+]^2 = (1.45) / (0.105)^2 = 131.5$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - (0.0592 / n) \log Q = 1.56 - (0.0592/2)\log(131.5)$$

$$E_{\text{cell}} = 1.56 - 0.063 = 1.50 \text{ v (after } \sim 90 \% \text{ completion)}$$



$$E^{\circ}_{\text{cell}} = -0.36 + 0.23 \text{ v} = -0.23 \text{ v}$$

$$\log K_{\text{sp}} = nE^{\circ}_{\text{cell}} / 0.0592 = 2 (-0.23) / (0.0592) = -7.77$$

$$K_{\text{sp}} = 1.7 \times 10^{-8}$$