CH 223 Chapter Seventeen Concept Guide

1. Half Reactions

Question

What are the half reactions for the reduction of iron(III) ion to iron(II) ion by tin(II) ion?

$$Sn^{2+} + 2\ Fe^{3+} \rightarrow 2\ Fe^{2+} + Sn^{4+}$$

Solution

Each iron(III) gains one electron during reduction and each tin(II) ion loses two electrons during oxidation.

Reduction half-reaction: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$

Oxidation half-reaction: $Sn^{2+}(aq) \rightarrow Sn^{4+} + 2e^-(aq)$

2. Ion-Electron Equations

Problem

Write the balanced half-reaction equation for the reduction of NO_3^- to $HNO_2(aq)$ in acidic solution.

Solution

Nitrate ion is reduced to nitrous acid: $NO_3^- \rightarrow HNO_2$

This equation is balanced with respect to the N atoms, but the oxygen atoms are unbalanced. Add one H_2O molecule to the products side to obtain

$$NO_3^- \rightarrow HNO_2 + H_2O$$

The H atoms are balanced in acidic solution by adding three H^+ atoms to the reactants side, giving

$$3\ H^+ + NO_3^- \rightarrow HNO_2 + H_2O$$

The charge is balanced by adding two e^- to the reactants side.

$$3\ H^+ + NO_3^- + 2\ e^- \rightarrow HNO_2 + H_2O$$

The total charge on the left is zero [(-1) + (3)(+1) + (2)(-1) = 0] and the total charge on the right is zero. The equation is balanced.

3. Cell Reactions

Problem

For a particular reaction, the anode, cathode, and overall cell reactions are:

$2\ H^+ + 2\ e^- \rightarrow H_2(g)$

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

$Mg(s) + 2\ H^+ \rightarrow Mg^{2+} + H_2(g)$

Label each reaction as the anode, cathode, and cell reaction.

Solution

The anode is the site of an oxidation reaction. The cathode is the site of a reduction reaction. The cell reaction is the result of the anode reaction added to the cathode reaction.

Cathode Reaction: $2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$

Anode Reaction: $Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

Cell Reaction: $Mg(s) + 2\ H^+ \rightarrow Mg^{2+}(aq) + H_2(g)$

4. Electrochemical Cells

Question

A galvanic cell is prepared by immersing a strip of magnesium metal into 1 M $Mg(NO_3)_2$ and a strip of copper metal into 1 M $CuSO_4$. The two solutions are separated by an unglazed porcelain wall.

(a) Write the anode reaction.

(b) Write the cathode reaction.

(c) Write the net cell reaction.

(d) From which electrode do the electrons flow to the other electrode if a wire is used to connect the two electrodes?

(e) What is the standard voltage of this cell (in Volts)?

Solution

(a) $Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

(b) $Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$

(c) $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

(d) The electrons flow from the magnesium electrode to the copper electrode if a wire connects the two electrodes.

(e) The voltage of this cell is 2.71 and is calculated by adding together the standard reduction potentials for the anode and cathode reaction.

$$\begin{array}{ll} Mg(s) \longrightarrow Mg^{2+}(aq) + 2\,e^- & -(-2.37) \\ Cu^{2+}(aq) + 2\,e^- \longrightarrow Cu(s) & 0.34 \\ \hline Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s) & \text{Cell Voltage} = 2.71 \end{array}$$

5. Thermodynamics of Electrochemical Cells

Question

What is the standard cell potential for an electrochemical cell in which the following spontaneous reaction takes place? $\Delta G°$ for this reaction is - 50.61 kJ.

$Cl_2(g) + 2\,Br^-(aq) \rightarrow Br_2(aq) + 2\,Cl^-(aq)$

Solution

To solve for E°, we need to use

$$E° = \frac{-\Delta G°}{nF}$$

Before solving for E°, we must find the value of *n*, the moles of electrons being transferred, by writing the half-reactions for the reduction of Cl_2 and oxidation of Br^-.

$$\begin{array}{l} Cl_2(g) + 2\,e^- \longrightarrow 2\,Cl^-(aq) \\ 2\,Br^-(aq) \longrightarrow Br_2(aq) + 2\,e^- \\ \hline Cl_2(g) + 2\,Br^-(aq) \longrightarrow Br_2(aq)\ \ 2\,Cl^-(aq) \end{array}$$

The value of *n* is 2.

The cell potential is found by solving for E° and substituting *n* = 2 and the known value of $\Delta G°$.

$E° = -(-50.61\text{ kJ})/(2\text{ mole }e^-)(96.5\text{ kJ/V mol }e^-) = 0.262\text{ V}$

6. Thermodynamics of Electrochemical Cells

Question

Under standard state conditions, the following reaction is spontaneous, and E° is 0.019 V.

$3\,Zn(s) + 2\,Cr^{3+}(aq) \rightarrow 3\,Zn^{2+}(aq) + 2\,Cr(s)$

Will this reaction occur spontaneously if $[Cr^{3+}] = 0.010$ mol/L and $[Zn^{2+}] = 5.5$ mol/L?

Solution

We will need to find the value of the reaction quotient and then solve for E.

$$Q = \frac{[Zn^{2+}]^3}{[Cr^{3+}]^2} = \frac{(5.5)^3}{(0.010)^2} = 1.7 \times 10^6$$

$$E = E^\circ - \left(\frac{0.0592}{n}\right)\left(\log Q\right) = 0.019 - \left(\frac{0.0592}{6}\right)\left(\log 1.7 \times 10^6\right)$$

$$E = -0.042$$

The negative cell potential implies that the reaction will not be spontaneous under these conditions. From Le Chatelier's principle, the reverse reaction would be expected to be favored.

7. Using the Faraday Constant

Question

A current of 1.65 amps is passed through a solution containing silver ions, Ag^+, for 10.0 minutes. The process causes silver to be deposited at the cathode. What mass of silver is deposited?

Solution

Step 1. Calculate the charge (number of coulombs) passed in 10.0 minutes.

Charge (Coulombs, C) = current (amps) x time (seconds)

$$\text{Charge (Coulombs, C)} = (1.65 \text{ amps})(10.0 \text{ minutes})\left(\frac{60 \text{ seconds}}{1 \text{ minute}}\right) = 990 \text{ C}$$

Step 2. Calculate the number of moles of electrons.

$$\left(990 \text{ C}\right)\left(\frac{1 \text{ mole e}^-}{9.65 \times 10^4 \text{ C}}\right) = 1.03 \times 10^{-2} \text{ mol e}^-$$

Step 3. Calculate the number of moles of silver and then the mass of silver deposited.

$$\left(1.03 \times 10^{-2} \text{ mol e}^-\right)\left(\frac{1 \text{ mol Ag}}{1 \text{ mole e}^-}\right)\left(\frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}}\right) = 1.11 \text{ g Ag}$$

8. Using the Faraday Constant

Question

A current of 2.00 amps is passed through a solution of $Ni(NO_3)_2$ for 2.00 hours. What mass of nickel is deposited at the cathode?

Solution

Step 1. Calculate the charge (number of coulombs) passed in 2 hours.

Charge (Coulombs, C) = current (amps) x time (seconds)

$$\text{Charge (Coulombs, C)} = (2.00\ \text{amps})(2.00\ \text{hours})\left(\frac{60\ \text{minutes}}{1\ \text{hour}}\right)\left(\frac{60\ \text{seconds}}{1\ \text{minute}}\right) = 1.44 \times 10^4\ \text{C}$$

Charge = 1.44×10^4 C

Step 2. Calculate the number of moles of electrons.

$$\left(1.44 \times 10^4\ \text{C}\right)\left(\frac{1\ \text{mole e}^-}{9.65 \times 10^4\ \text{C}}\right) = 0.149\ \text{mol e}^-$$

Step 3. Calculate the number of moles of nickel and then the mass of nickel deposited.

$$\left(0.149\ \text{mol e}^-\right)\left(\frac{1\ \text{mol Ni}}{2\ \text{mole e}^-}\right)\left(\frac{58.693\ \text{g Ni}}{1\ \text{mol Ni}}\right) = 4.37\ \text{g Ni}$$

9. Batteries

Question

In this lesson, the anode, cathode, and net reactions for dry cell and alkaline batteries are shown. Based on these reactions, what is a possible disadvantage to a dry cell battery compared to alkaline battery?

Solution

In the dry cell battery, the products of the cathode reaction are gases. If current is drawn from the battery rapidly, the gaseous products cannot be consumed rapidly enough, resulting in a drop in potential. In the alkaline battery, no gases are formed, thus there is no decline in potential under high current loads.

10. Batteries

Problem

In principle, a battery can be made from aluminum metal and chlorine gas. Write a balanced equation for the reaction that occurs in a battery using $Al^{3+}(aq)/Al(s)$ and $Cl_2(g)/Cl^-(aq)$ half-reactions, and indicate which half-reaction occurs at the anode and which at the cathode.

Solution

$$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ Al^{3+}(aq) + 6\ Cl^-(aq)$$

Al is the anode and the Cl_2/Cl^- compartment is the cathode.

11. Corrosion

Problem

A 5.0-kg magnesium bar is attached to a buried iron pipe to protect the pipe from corrosion. Explain how the magnesium protects the pipe.

Solution

E° for magnesium oxidation is 2.37 V. E° for iron oxidation is only 0.44 V. Magnesium is the better oxidizing agent, and thus protects the iron pipe because the metal oxidizes more readily than iron. It is a sacrificial reductant.