Chapter 13 Cells and Batteries

Solutions to Practice Problems

1.
Problem
(**Note:** Obtain the necessary standard reduction potential values from Table 13.1 or the table in the Appendix.) Write the two half-reactions for the following redox reaction. Find the standard cell potential for a voltaic cell in which this reaction occurs.

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

What Is Required?
You must write the half-reactions for the redox reaction and determine the cell potential for the reaction.

What Is Given?
The balanced net ionic equation and a table of standard reduction potentials are given.

Plan Your Strategy
Write the oxidation and reduction half-reactions.
Locate the relevant reduction potentials in a table of standard reduction potentials.
Subtract the reduction potentials to find the cell potential by using the formula

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

Act on Your Strategy
Oxidation half-reaction (occurs at the anode): $2Br^-(aq) \rightarrow Br_2(l) + 2e^-$
Reduction half-reaction (occurs at the cathode): $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
The standard reduction potentials from the Chemistry Data Booklet are as follows:
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ $E^o_{cathode} = +1.36$ V
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$ $E^o_{anode} = +1.07$ V

The cell potential is
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 1.36\text{ V} - 1.07\text{ V} = +0.29\text{ V}$

Check Your Solution
The E°_{cell} is a positive value indicating that the reaction occurs spontaneously as written and that $Br^-(aq)$ is a stronger reducing agent than $Cl^-(aq)$ which is consistent with the standard reduction potentials. The answer has the correct units (V) and the correct number of digits after the decimal (2).

2.
Problem

(**Note:** Obtain the necessary standard reduction potential values from Table 13.1 or the table in the Appendix.) Write the two half-reactions for the following redox reaction. Find the standard cell potential for a voltaic cell in which this reaction occurs.
$Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$

What is Required?
You must write the two half-reactions for the redox reaction and calculate the standard cell potential.

What is Given?
The balanced redox reaction is given.

Plan Your Strategy
Write the half reactions.
Find the standard reduction potential values from Table 13.1.
Calculate the standard cell potential using the equation $E°_{cell} = E°_{cathode} - E°_{anode}$.

Act on Your Strategy
$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E°_{cathode} = 0.80$ V
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ $E°_{anode} = –2.37$ V

$E°_{cell} = E°_{cathode} - E°_{anode} = –0.80 – (–2.37\text{ V}) = –3.17$ V

Check Your Solution
The answer seems reasonable and has the correct units (V).

3.
Problem
(**Note:** Obtain the necessary standard reduction potential values from Table 13.1 or the table in the Appendix.) Write the two half-reactions for the following redox reaction. Find the standard cell potential for a voltaic cell in which this reaction occurs.
$Sn(s) + 2HBr(aq) \rightarrow SnBr_2(aq) + H_2(g)$

What Is Required?
You must write the half-reactions for the redox reaction and determine the cell potential for the reaction.

What Is Given?
The balanced equation and a table of standard reduction potentials are given.

Plan Your Strategy
Write the total ionic equation for the reaction. Cancel ions common to both sides of the equation and write the net ionic equation.
Write the oxidation and reduction half-reactions.
Locate the relevant reduction potentials in a table of standard reduction potentials.
Subtract the reduction potentials to find the cell potential by using the formula

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

Act on Your Strategy

Total ionic equation: $Sn(s) + 2H^+(aq) + 2Br^-(aq) \rightarrow Sn^{2+}(aq) + 2Br^-(aq) + H_2(g)$
Net ionic equation: $Sn(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$
Oxidation half-reaction (occurs at the anode): $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$
Reduction half-reaction (occurs at the cathode): $2H^+(aq) + 2e^- \rightarrow H_2(g)$
The standard reduction potentials from the Chemistry Data Booklet are as follows:

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $\quad E^o_{cathode} = +0.00$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ $\quad E^o_{anode} = -0.14$ V

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.00\text{ V} - (-0.14\text{ V}) = +0.14\text{ V}$

Check Your Solution

The E°_{cell} is a positive value indicating that the reaction occurs spontaneously as written and that Sn(s) is a stronger reducing agent than $H_2(g)$ which is consistent with the standard reduction potentials. The answer has the correct units (V) and the correct number of digits after the decimal (2).

4.

Problem

(**Note:** Obtain the necessary standard reduction potential values from Table 13.1 or the table in the Appendix.) Write the two half-reactions for the following redox reaction. Find the standard cell potential for a voltaic cell in which this reaction occurs.
$Cr(s) + 3AgCl(s) \rightarrow CrCl_3(aq) + 3Ag(s)$

What Is Required?

You must write the half-reactions for the redox reaction and determine the cell potential for the reaction.

What Is Given?

The balanced equation and a table of standard reduction potentials are given.

Plan Your Strategy

Write the total ionic equation for the reaction. Cancel ions common to both sides of the equation and write the net ionic equation.
Write the oxidation and reduction half-reactions.
Locate the relevant reduction potentials in a table of standard reduction potentials.
Subtract the reduction potentials to find the cell potential by using the formula

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

Act on Your Strategy

Net ionic equation: $Cr(s) + 3AgCl(s) \rightarrow Cr^{3+}(aq) + 3Cl^-(aq) + 3Ag(s)$

Oxidation half-reaction (occurs at the anode): $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$
Reduction half-reaction (occurs at the cathode): $AgCl(s) \rightarrow Ag(s) + Cl^-(aq)$
The standard reduction potentials from the Chemistry Data Booklet are as follows:
$AgCl(s) + 1e^- \rightarrow Ag(s) + Cl^-(aq)$ $\quad E^o_{cathode} = +0.22\text{ V}$
$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ $\quad E^o_{anode} = -0.74\text{ V}$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.22\text{ V} - (-0.74\text{ V}) = +0.96\text{ V}$

Check Your Solution
The E°_{cell} is a positive value indicating that the reaction occurs spontaneously as written and that Cr(s) is a stronger reducing agent than Ag(s) which is consistent with the standard reduction potentials. The answer has the correct units (V) and the correct number of digits after the decimal (2).

5.
Problem
One half-cell of a voltaic cell has a nickel electrode in a 1 mol/L nickel(II) chloride solution. The other half-cell has a cadmium electrode in a 1 mol/L cadmium chloride solution.
a) Find the cell potential.
b) Identify the anode and the cathode.
c) Write the oxidation half-reaction, the reduction half-reaction, and the overall cell reaction.

What is Required?
a) You must calculate the cell potential for a Ni-Cd voltaic cell.
b) You must identify the anode and cathode in the Ni-Cd voltaic cell.
c) You must write the oxidation half-reaction, the reduction half-reaction, and the overall cell reaction for the Ni-Cd voltaic cell

What is Given?
The Ni electrode is in 1.0 mol/L $NiCl_2(aq)$ and the Cd electrode is in 1.0 mol/L $CdCl_2(aq)$. The two electrodes are joined to form a voltaic cell.

Plan Your Strategy
Locate the relevant reduction potentials in a table of standard reduction potentials for the reduction of $Ni^{2+}(aq)$ and $Cd^{2+}(aq)$.
Write the oxidation and reduction half-reactions.
Subtract the reduction potentials to find the cell potential by using the formula

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

Act on Your Strategy
The standard reduction potentials from the Chemistry Data Booklet are as follows:
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $\quad E^o_{cathode} = -0.26\text{ V}$
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$ $\quad E^o_{anode} = -0.40\text{ V}$

a) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.26\ V - (-0.40\ V) = +0.14\ V$

b) The anode is the electrode where oxidation occurs. This is the Cd electrode.
The cathode is the electrode where reduction occurs. This is the Ni electrode.

c) Oxidation: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$
Reduction: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$
Overall cell reaction: $Cd(s) + Ni^{2+}(aq) \rightarrow Ni(s) + Cd^{2+}(aq)$

Check Your Solution
A positive E°_{cell} is obtained indicating that the reaction occurs spontaneously. The answer has the correct units (V) and the correct number of digits after the decimal (2).

6.
Problem
An external voltage is applied to change the voltaic cell in question 5 into an electrolytic cell. Repeat parts (a) to (c) for the electrolytic cell.

What is Required?
a) You must calculate the cell potential for a Ni-Cd electrolytic cell.
b) You must identify the anode and cathode in the Ni-Cd electrolytic cell.
c) You must write the oxidation half-reaction, the reduction half-reaction, and the overall cell reaction for the Ni-Cd electrolytic cell

What is Given?
The Ni electrode is in 1.0 mol/L $NiCl_2(aq)$ and the Cd electrode is in 1.0 mol/L $CdCl_2(aq)$. An external voltage is applied to operate this combination of electrodes as an electrolytic cell.

Plan Your Strategy
Locate the relevant reduction potentials in a table of standard reduction potentials for the reduction of $Ni^{2+}(aq)$ and $Cd^{2+}(aq)$. The external voltage will force the spontaneous reaction in the reverse direction. Therefore the reduction half-reaction having the more positive E^o will be forced to occur as an oxidation and will be the anode and the reduction half reaction with the less positive E^o will occur as the reduction and will be the cathode.
Write the oxidation and reduction half-reactions.
Subtract the reduction potentials to find the cell potential by using the formula

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

Act on Your Strategy
The standard reduction potentials from the Chemistry Data Booklet are as follows:
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $E^o_{anode} = -0.26\ V$
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$ $E^o_{cathode} = -0.40\ V$

a) $E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode} = -0.40\text{ V} - (-0.26\text{ V}) = -0.14\text{ V}$

b) The anode is the electrode where oxidation occurs. This is the Ni electrode.
The cathode is the electrode where reduction occurs. This is the Cd electrode.

c) Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^{-}$
Reduction: $Cd^{2+}(aq) + 2e^{-} \rightarrow Cd(s)$
Overall cell reaction: $Cd^{2+}(aq) + Ni(s) \rightarrow Ni^{2+}(aq) + Cd(s)$

7.
Problem
Predict the products of the electrolysis of a 1 mol/L solution of cobalt chloride.

What Is Required?
You need to predict the products of the electrolysis of 1.0 mol/L $CoCl_2(aq)$.

What Is Given?
The name and concentration of the aqueous solution are given and a table of standard reduction potentials is available. The non-standard reduction potentials for water are known.

Plan Your Strategy
List the possible half-reactions that could occur at each electrode with their reduction potentials. Since the $Co^{2+}(aq)$ and $Cl^{-}(aq)$ concentrations are 1.0 mol/L, the standard reduction potentials for the half-reactions that involve these ions can be used. Use the non-standard values for water. Take into consideration a 0.6-V overpotential for the half-reactions involving water. Select the oxidation and reduction half-reactions that require the least amount of energy. These are the reactions that have the smaller negative reduction potential. Combine the selected half-reactions.

Act on Your Strategy
The two possible oxidation half-reactions at the anode are the oxidation of chloride ion or the oxidation of water:
$2Cl^{-}(aq) \rightarrow Cl_2(g) + 2e^{-}$ $\quad E^{o}_{anode} = -1.36\text{ V}$
$2H_2O(l) \rightarrow O_2(g) + 4H^{+}(aq) + 4e^{-}$ $\quad -E^{o}_{anode} = -0.82\text{ V}$ with overpotential $= -1.42\text{ V}$

The two possible reduction half-reactions at the cathode are the reduction of cobalt ions or the reduction of water:
$Co^{2+}(aq) + 2e^{-} \rightarrow Co(s)$ $\quad E^{o}_{cathode} = -0.28\text{ V}$
$2H_2O(l) + 2e^{-} \rightarrow H_2(g) + 2OH^{-}(aq)$ $\quad E^{o}_{cathode} = -0.42\text{ V}$ with overpotential $= -1.02\text{ V}$

At the anode, the oxidation of the chloride has the smaller negative reduction potential and at the cathode the reduction of cobalt has the smaller negative reduction potential. The combined half-reactions are
$Co^{2+}(aq) + 2Cl^{-}(aq) \rightarrow Co(s) + Cl_2(g)$
The products of the electrolysis of 1.0 mo/L $CoCl_2(aq)$ are Co(s) and $Cl_2(g)$.

Check Your Solution
The predicted cell voltage would be
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
$= -0.28\text{ V} - 1.36\text{ V} = -1.64\text{ V}$
The negative cell voltage is expected in an electrolysis reaction and the products are reasonable under these conditions.

8.
Problem
Predict the products of the electrolysis of a 1 mol/L solution of silver nitrate.

What Is Required?
You need to predict the products of the electrolysis of 1.0 mol/L $AgNO_3$(aq).

What Is Given?
The name and concentration of the aqueous solution are given and a table of standard reduction potentials is available. The non-standard reduction potentials for water are known.

Plan Your Strategy
List the possible half-reactions that could occur at each electrode with their reduction potentials. Since the Ag^+(aq) concentrations is 1.0 mol/L, the standard reduction potentials for the half-reactions for this ion can be used.
Use the non-standard reduction potentials for water and take into consideration a 0.6-V overpotential for the half-reactions involving water. Select the oxidation and reduction half-reactions that require the least amount of energy. These are the reactions that have the smaller negative reduction potential. Combine the selected half-reactions.

Act on Your Strategy
The only possible oxidation half-reaction that can occur at the anode is the oxidation of water. This is because the nitrogen in NO_3^-(aq) cannot be oxidized any further.

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ $\quad -E_{anode} = -0.82\text{ V}$ with overpotential $= -1.42\text{ V}$

The two possible reduction half-reactions at the cathode are the reduction of silver ions or the reduction of water:
$Ag^+(aq) + e^- \rightarrow Ag(s)$ $\quad E^o_{cathode} = +\ 0.80\text{ V}$
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $\quad E^o_{cathode} = -0.42\text{ V}$ with overpotential $= -1.02\text{ V}$

At the anode, the product is O_2(g). At the cathode the reduction of silver has the smaller negative (more positive) reduction potential.
The combined half-reactions are
$4Ag^+(aq) + 2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4Ag(s)$

The products of the electrolysis of 1.0 mo/L $AgNO_3$(aq) are Ag(s) and O_2(g).

Check Your Solution
The predicted cell voltage would be
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
$= +0.80 \text{ V} - 1.42 \text{ V} = -0.62 \text{ V}$
The negative cell voltage is expected in an electrolysis reaction and the products are reasonable under these conditions.

9.
Problem
Calculate the mass of zinc plated onto the cathode of an electrolytic cell by a current of 750 mA in 3.25 h.

What Is Required?
You must calculate the mass of zinc plated onto the cathode.

What Is Given?
The element plated is zinc.
Current = 750 mA = 0.750 A
Time = 3.25 h = 11 700 s
Charge on one mole of electrons is 96 500 C/mol.
MZn(s) = 65.39 g/mol

Plan Your Strategy
Calculate the quantity of electricity using: charge (C) = current (A) × time (s)
Using the relationship that the charge on 1 mole of electrons is 96 500 C/mol, calculate the number of moles of electrons that pass through the circuit.
Use the half-reaction $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ to calculate the number of moles of Zn(s) produced.
Use the molar mass of zinc to convert the amount of zinc to a mass: $m = n \times M$

Act on Your Strategy
Quantity of electricity (C) = current (A) × time (s)
$= 0.750 \text{ A} \times 1.17 \times 10^4 \text{ s}$
$= 8.775 \times 10^3 \text{ C}$

$$\text{Amount of electrons} = 8.775 \times 10^3 \text{ C} \times \frac{1 \text{ mol e}^-}{96500 \text{ C}} = 9.093 \times 10^{-2} \text{ mol e}^-$$

$$n\text{Zn(s) plated} = 9.093 \times 10^{-2} \text{ mol e}^- \times \frac{1 \text{ mol Zn}}{2 \text{e}^-} = 4.547 \times 10^{-2} \text{ mol e}^-$$

$$m\text{Zn(s)} = n \times M = 4.547 \times 10^{-2} \text{ mol e}^- \times 65.39 \text{ g/mol} = 2.97 \text{ g}$$

Check Your Solution
The answer is reasonable for this data, has the correct units (g) and the correct number of significant digits (3).

10.
Problem
How many minutes does it take to plate 0.925 g of silver onto the cathode of an electrolytic cell using a current of 1.55 A?

What Is Required?
You must calculate the time, in minutes, to plate the given mass of silver onto the cathode of an electrolytic cell.

What Is Given?
mass of silver plated = m = 0.925 g
current passing through the cell = 1.55 A
charge on one mole of electrons = 96 500 C/mol
MAg(s) = 107.87 g/mol

Plan Your Strategy
Convert the mass of silver to moles: $n = \frac{m}{M}$

Use the half-reaction that occurs at the cathode, $Ag^{+}(aq) + 1\ e^{-} \rightarrow Ag(s)$, to calculate the number of moles of electrons that pass through the circuit. Convert this number of moles of electrons to a quantity of electricity using the relationship that the charge on 1 mole of electrons is 96 500 C/mol. Solve for the time that the current flows using:
Quantity of electricity that passes through the circuit = current (A) × time (s)
Convert the time from seconds to minutes.

Act on Your Strategy
$$n\text{Ag(s)} = 0.925\ \text{g Ag} \times \frac{1\ \text{mol}}{107.87\text{g}} = 8.575 \times 10^{-3}\ \text{mol}$$

$$\text{number of moles of electrons} = 8.575 \times 10^{-3}\ \text{mol Ag(s)} \times \frac{1\ \text{mol e}^{-}}{1\ \text{mol Ag(s)}}$$
$$= 8.575 \times 10^{-3}\ \text{mol e}^{-}$$

$$\text{Quantity of electricity} = 8.575 \times 10^{-3}\ \text{mol e}^{-} \times 96\ 500\ \text{C/mol e}^{-}$$
$$= 827.5\text{C}$$

$$\text{time (s)} = \frac{\text{quantity of electricity(C)}}{\text{current (A)}} = \frac{827.5\text{C}}{1.55\text{A}} = 533.9\ \text{s}$$

$533.9 \text{ s} \times \dfrac{1\,\text{min}}{60\,\text{s}} = 8.90 \text{ min}$

Check Your Solution
The answer is reasonable for this data and has the correct units (min) and the correct number of significant digits (3).

11.
Problem
The nickel anode in an electrolytic cell decreases in mass by 1.20 g in 35.5 min. The oxidation half-reaction converts nickel atoms to nickel(II) ions. What is the average current?

What Is Required?
You must calculate the average current required to remove the given mass of nickel from the anode in the given length of time.

What Is Given?
mass of Ni(s) removed = m = 1.20 g
time = 35.5 min = 2130 s
half reaction for the oxidation of nickel is $Ni^{2+}(aq) + 2e^{-} \rightarrow Ni(s)$
charge on one mole of electrons = 96 500 C/mol
MNi(s) = 58.69 g/mol

Plan Your Strategy
Convert the mass of nickel to moles: $n = \dfrac{m}{M}$

Use the oxidation half reaction to calculate the number of moles of electrons that pass through the circuit.
Calculate the quantity of electricity that passes through the circuit using the relationship that the charge on 1 mol of electrons is 96 500 C/mol.
Determine the current that flows in the circuit using the equation:
Quantity of electricity (C) = current × time (s)

Act on Your Strategy
$n\text{Ni(s)} = 1.20 \text{ g Ni(s)} \times \dfrac{1\,\text{mol}}{58.69\,\text{g}} = 2.045\times 10^{-2} \text{ mol}$

$$\text{number of moles of electrons} = 2.045\times 10^{-2} \text{ mol Ni(s)} \times \frac{2\,\text{mol e}^{-}}{1\,\text{mol Ni(s)}}$$
$$= 4.089 \times 10^{-2} \text{ mol e}^{-}$$

Quantity of electricity $= 4.089 \times 10^{-2}$ mol $e^{-} \times 96\,500$ C/mol e^{-}
$= 3946$ C

$$\text{current(A)} = \frac{\text{quantity of electricity(C)}}{\text{time(s)}} = \frac{3946\,\text{C}}{2130\,\text{s}} = 1.85\ \text{A}$$

The average current that flows is 1.85 A

Check Your Solution
The answer is reasonable for this data and has the correct unit (A) and the correct number of significant digits (3).

12.
Problem
The following two half-reactions take place in an electrolytic cell with an iron anode and a chromium cathode:
Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$
Reduction: $Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$
During the process, the mass of the iron anode decreases by 1.75 g.
a) Find the change in mass of the chromium cathode.
b) Explain why you do not need to know the electric current or the time to complete part (a).

What is Required?
You must calculate the change in mass that occurs at the chromium electrode.

What is Given?
The two half reactions are:
Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$
Reduction: $Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$
change in mass at iron electrode = m = 1.75 g
MFe(s) = 55.85 g/mol
MCr(s) = 52.00 g/mol

Plan Your Strategy
Write the balance net ionic equation for the overall reaction that occurs.

Convert the mass of iron to moles: $n = \dfrac{m}{M}$

Use the balanced overall equation to calculate the number of moles of chromium that is produced in the reaction. Calculate the mass of chromium: $m = n \times M$

Act on Your Strategy

a) Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$

Reduction: $Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$

The LCM of 2 and 3 is 6.

Multiply the oxidation half-reaction by 3, and the reduction half-reaction by 2 so that equal numbers of electrons are lost and gained and add these two half reactions.

$$3Fe(s) \rightarrow 3Fe^{2+}(aq) + 6e^{-}$$
$$2Cr^{3+}(aq) + 6e^{-} \rightarrow 2Cr(s)$$

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

The balanced net ionic equation shows that 2 mol of chromium are formed for every 3 mol of iron that react.

$$n\text{Fe(s)} = \frac{m}{M} = \frac{1.75\,\text{g}}{55.85\,\text{g/mol}} = 0.03133 \text{ mol Fe(s)}$$

Use mole ratio from the balanced equation to calculate the number of moles of chromium.

$$0.03133 \text{ mol Fe(s)} \times \frac{2\,\text{mol Cr}}{3\,\text{mol Fe}} = 0.02089 \text{ mol Cr(s)}$$

mCr(s) = 0.02089 mol Cr × 52.00 g/mol = 1.09 g Cr(s)

The chromium cathode increases in mass by 1.09 g

b) Since the overall balanced equation can be determined from the given information, the mole ratio is sufficient information to calculate the mass change.

Check Your Solution

The answer is reasonable for this data and has the correct units (g) and the correct number of significant digits (3).