B11 · Faraday's law with the unknown chosen

Completed notes

Goals: Choose the starting point of the Faraday chain from the givens and the question. Find a current, a time, an ion charge or a mass in an electrolysis. Use Faraday’s law in a voltaic cell and use the mass gained at a cathode.

Slide 1 of 8 · One chain, any unknown

Every Faraday problem uses the same chain: current and time, charge, moles of **electrons**, moles of metal, and **mass**.

Step one: write the **givens** and the question. Step two: start at a given quantity and walk toward the **unknown**.

Starting with current and time means going forward: multiply for charge, **divide** by F, then divide by the electron coefficient.

Starting with a mass means going backward: **multiply** by the electron coefficient, then multiply by **F** to get charge.

Example. A problem gives the mass of silver deposited and the time, and asks for the current. Which quantity do you start with, which direction do you walk, and what do you do at each step?

1. Start with the mass, so walk backward along the chain.

2. $n(Ag) = m \div M$, then multiply by the coefficient of e^- to get $n(e^-)$.

3. Multiply $n(e^-)$ by F to get the charge q.

4. Change the time to seconds, then $I = q \div t$.

Try it. A problem gives the current and the time and asks for the mass of copper deposited. Which direction do you walk, and what do you do at each step?

1. Start with current and time, so walk forward along the chain.

2. $q = It$, with t in seconds.

3. $n(e^-) = q \div F$, then divide by 2 for $Cu^{2+} + 2e^- \rightarrow Cu$ to get n(Cu).

4. $m = nM$.

Slide 2 of 8 · Finding the current

The question gives a mass and a time and asks for the current. Start with the mass and go **backward**: n(Cu) = m ÷ M.

In $Cu^{2+} + 2e^- \rightarrow Cu$, multiply the moles of copper by **2** to get moles of electrons. Then multiply by **F** to get the charge.

Current is charge divided by time: I = **q ÷ t**. The time must be in **seconds**.

Leaving the time in minutes makes the current **60 times** too big.

Example. A copper cathode gains 1.25 g in 40.0 minutes from Cu^{2+} solution. What constant current was used? M(Cu) = 63.55 g/mol.

1. n(Cu) = 1.25 ÷ 63.55 = 0.01967 mol

2. $Cu^{2+} + 2e^- \rightarrow Cu$, so $n(e^-) = 2 \times 0.01967 = 0.03934$ mol

3. $q = 0.03934 \times (9.65 \times 10^4) = 3796$ C

4. Time in seconds: 40.0 × 60 = 2400 s

5. I = q ÷ t = 3796 ÷ 2400 = 1.58 A

Try it. A zinc cathode gains 0.450 g in 30.0 minutes from Zn^{2+} solution. What constant current was used? M(Zn) = 65.41 g/mol.

1. n(Zn) = 0.450 ÷ 65.41 = 0.006880 mol

2. $Zn^{2+} + 2e^- \rightarrow Zn$, so $n(e^-) = 2 \times 0.006880 = 0.01376$ mol

3. $q = 0.01376 \times (9.65 \times 10^4) = 1328$ C

4. Time in seconds: 30.0 × 60 = 1800 s

5. I = q ÷ t = 1328 ÷ 1800 = 0.738 A

Slide 3 of 8 · Finding the ion charge

When the half-reaction is unknown, let **z** be the number of electrons each ion gains. It is also the **charge** on the ion.

Find the charge first: q = It. Then the moles of electrons are n(e^-) = **q ÷ F**.

Find the moles of metal from its mass: n(metal) = **m ÷ M**.

Then z = n(e^-) ÷ n(metal). Round z to the nearest **whole number**. If z = 2.00, the ion has a charge of **2+**.

Example. A metal with a molar mass of 63.55 g/mol is plated from a solution of its ions by 2.00 A for 20.0 minutes. The mass deposited is 0.790 g. What is the charge on the ion?

1. Time in seconds: 20.0 × 60 = 1200 s

2. q = It = 2.00 × 1200 = 2400 C

3. n(e^-) = 2400 ÷ (9.65 × 10^4) = 0.02487 mol

4. n(metal) = 0.790 ÷ 63.55 = 0.01243 mol

5. z = n(e^-) ÷ n(metal) = 0.02487 ÷ 0.01243 = 2.00

6. The ion charge is 2+, so the half-reaction is $M^{2+} + 2e^- \rightarrow M$.

Try it. A metal with a molar mass of 58.69 g/mol is plated from a solution of its ions by 1.00 A for 40.0 minutes. The mass deposited is 0.730 g. What is the charge on the ion?

1. Time in seconds: 40.0 × 60 = 2400 s

2. q = It = 1.00 × 2400 = 2400 C

3. n(e^-) = 2400 ÷ (9.65 × 10^4) = 0.02487 mol

4. n(metal) = 0.730 ÷ 58.69 = 0.01244 mol

5. z = 0.02487 ÷ 0.01244 = 2.00

6. The ion charge is 2+.

Slide 4 of 8 · The chain in a voltaic cell

In a voltaic cell the electrons in the wire come from the **anode** half-reaction and are used in the **cathode** half-reaction.

The charge that flows is q = It, so n(e⁻) = **q ÷ F**. These same electrons pass through both electrodes.

At the anode, Zn → Zn^{2+} + **2e⁻**, so n(Zn) = n(e⁻) ÷ 2. The zinc anode **loses** mass.

At the cathode, Cu^{2+} + 2e⁻ → Cu, so the copper cathode **gains** mass. The moles are equal, but the masses differ because the **molar masses** differ.

Example. A zinc and copper voltaic cell gives a steady 0.500 A for 30.0 minutes. What mass of zinc dissolves from the anode, and what mass of copper deposits on the cathode? M(Zn) = 65.41 g/mol and M(Cu) = 63.55 g/mol.

1. Time in seconds: 30.0 × 60 = 1800 s

2. q = It = 0.500 × 1800 = 900 C

3. n(e⁻) = 900 ÷ (9.65 × 10^4) = 0.009326 mol

4. Anode: Zn → Zn^{2+} + 2e⁻, so n(Zn) = 0.009326 ÷ 2 = 0.004663 mol

5. Zinc lost: 0.004663 × 65.41 = 0.305 g

6. Cathode: Cu^{2+} + 2e⁻ → Cu, so n(Cu) = 0.004663 mol

7. Copper gained: 0.004663 × 63.55 = 0.296 g

Try it. A zinc and copper voltaic cell gives a steady 0.250 A for 2.00 hours. What mass of zinc dissolves from the anode, and what mass of copper deposits on the cathode? M(Zn) = 65.41 g/mol and M(Cu) = 63.55 g/mol.

1. Time in seconds: 2.00 × 3600 = 7200 s

2. q = It = 0.250 × 7200 = 1800 C

3. n(e⁻) = 1800 ÷ (9.65 × 10^4) = 0.01865 mol

4. n(Zn) = n(Cu) = 0.01865 ÷ 2 = 0.009325 mol

5. Zinc lost: 0.009325 × 65.41 = 0.610 g

6. Copper gained: 0.009325 × 63.55 = 0.593 g

Slide 5 of 8 · Your turn: gases from an aqueous solution

The same current supplies electrons at **both** electrodes, so one value of $n(e^-)$ feeds both half-reactions.

Cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$. One H_2 needs **2** electrons, so $n(H_2) = n(e^-) \div 2$.

Anode: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$. One O_2 releases **4** electrons, so $n(O_2) = n(e^-) \div$ **4**.

Check: $n(H_2)$ is **twice** $n(O_2)$, which matches $2H_2O \rightarrow 2H_2 + O_2$. Read whether the question asks for **moles** or grams.

Example. $Na_2SO_4(aq)$ is electrolyzed with inert electrodes using 2.00 A for 30.0 minutes. The half-reactions are $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ and $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$. Find the moles of hydrogen and oxygen formed.

1. Time in seconds: $30.0 \times 60 = 1800$ s

2. $q = It = 2.00 \times 1800 = 3600$ C

3. $n(e^-) = 3600 \div (9.65 \times 10^4) = 0.037306$ mol

4. $n(H_2) = 0.037306 \div 2 = 0.018653$ mol, which rounds to 0.0187 mol

5. $n(O_2) = 0.037306 \div 4 = 0.0093265$ mol, which rounds to 0.00933 mol

6. Check: 0.0187 is about twice 0.00933, which matches the 2 to 1 ratio.

Try it. $Na_2SO_4(aq)$ is electrolyzed with inert electrodes using 1.60 A for 30.0 minutes. Find the moles of hydrogen and oxygen formed.

1. Time in seconds: $30.0 \times 60 = 1800$ s

2. $q = It = 1.60 \times 1800 = 2880$ C

3. $n(e^-) = 2880 \div (9.65 \times 10^4) = 0.02984$ mol

4. $n(H_2) = 0.02984 \div 2 = 0.0149$ mol

5. $n(O_2) = 0.02984 \div 4 = 0.00746$ mol

Slide 6 of 8 · Mass gained, then time

When an electrode is weighed twice, subtract: mass gained = **final mass − starting mass**. Do not use the **final mass** alone.

Then n(Cu) = m ÷ M with M = **63.55 g/mol**. Because $Cu^{2+} + 2e^- \rightarrow Cu$, $n(e^-) = \mathbf{2} \times n(Cu)$.

The charge is $q = n(e^-) \times \mathbf{F}$. The time is $t = q \div \mathbf{I}$, where I is the current in amperes.

Give the time in the unit the question asks for. Divide seconds by **60** to get minutes.

Example. A cathode has a mass of 14.200 g before copper plating and 14.850 g after. The current was 1.50 A. How long did the plating run, in minutes? M(Cu) = 63.55 g/mol.

1. Mass gained = 14.850 − 14.200 = 0.650 g

2. n(Cu) = 0.650 ÷ 63.55 = 0.01023 mol

3. $n(e^-) = 2 \times 0.01023 = 0.02046$ mol

4. $q = 0.02046 \times (9.65 \times 10^4) = 1974$ C

5. t = q ÷ I = 1974 ÷ 1.50 = 1316 s

6. t = 1316 ÷ 60 = 21.9 min

Try it. A cathode has a mass of 25.300 g before silver plating and 26.085 g after. The current was 0.900 A. How long did the plating run, in minutes? The half-reaction is $Ag^+ + e^- \rightarrow Ag$. M(Ag) = 107.87 g/mol.

1. Mass gained = 26.085 − 25.300 = 0.785 g

2. n(Ag) = 0.785 ÷ 107.87 = 0.007277 mol

3. Ratio 1 to 1, so $n(e^-) = 0.007277$ mol

4. $q = 0.007277 \times (9.65 \times 10^4) = 702.2$ C

5. t = q ÷ I = 702.2 ÷ 0.900 = 780.2 s

6. t = 780.2 ÷ 60 = 13.0 min

Slide 7 of 8 · Method: how to solve a Faraday problem when the unknown changes

Step 1. Write the **chain**: I and t, then q, then n(e^-), then n(substance), then m.

Step 2. Mark the **given** piece and the unknown piece.

Step 3. Move from the given toward the unknown. **Flip** each step when you go backward.

Step 4. **Solve**. Carry an extra digit in the middle lines.

Step 5. Run the chain **forward** with your answer and compare to the given.

Example. Use the method on a cathode that has a mass of 14.200 g before copper plating and 14.850 g after. The current was 1.50 A. Find the time.

1. Step 2: the mass gained is 14.850 − 14.200 = 0.650 g. The unknown is t.

2. Step 3: n(Cu) = 0.650 ÷ 63.55 = 0.01023 mol. n(e^-) = 2 × 0.01023 = 0.02046 mol. q = 0.02046 × (9.65×10^4) = 1974 C. t = 1974 ÷ 1.50 = 1316 s.

3. Step 4: t = 1316 s ÷ 60 = 21.9 min.

4. Step 5: forward, 1.50 × 1316 = 1974 C, which gives back 0.650 g.

Try it. Use the method on a cathode that gains 0.850 g of silver in 15.0 minutes from $Ag^+ + e^- \rightarrow Ag$. M(Ag) = 107.87 g/mol. Find the current.

1. t = 15.0 × 60 = 900 s. n(Ag) = 0.850 ÷ 107.87 = 0.00788 mol.

2. n(e^-) = 0.00788 mol. q = 0.00788 × (9.65×10^4) = 760 C.

3. I = 760 ÷ 900 = 0.844 A.

Slide 8 of 8 · Mass at the anode

In copper refining, the impure copper is the **anode**. It dissolves: $Cu \rightarrow Cu^{2+} + \mathbf{2e^-}$.

The same electrons leave the anode and arrive at the **cathode**, so the Faraday chain works at both electrodes.

Change hours to seconds first: 2.00 h = **7200** s. Then $q = It$ and $n(e^-) = q \div F$.

Each copper atom gives **2** electrons, so $n(Cu) = n(e^-) \div 2$. The anode **loses** this mass.

Example. In copper refining, a copper anode dissolves and pure copper plates onto the cathode. A current of 10.0 A runs for 2.00 hours. What mass of copper dissolves from the anode? M(Cu) = 63.55 g/mol.

1. Time in seconds: $2.00 \times 3600 = 7200$ s

2. $q = It = 10.0 \times 7200 = 7.20 \times 10^4$ C

3. $n(e^-) = 7.20 \times 10^4 \div (9.65 \times 10^4) = 0.746$ mol

4. Anode: $Cu \rightarrow Cu^{2+} + 2e^-$, so $n(Cu) = 0.746 \div 2 = 0.373$ mol

5. $m = nM = 0.373 \times 63.55 = 23.7$ g. The anode loses 23.7 g.

Try it. In copper refining, a current of 6.00 A runs for 1.00 hour. What mass of copper dissolves from the anode? M(Cu) = 63.55 g/mol.

1. Time in seconds: $1.00 \times 3600 = 3600$ s

2. $q = It = 6.00 \times 3600 = 2.16 \times 10^4$ C

3. $n(e^-) = 21\,600 \div (9.65 \times 10^4) = 0.2238$ mol

4. Anode: $Cu \rightarrow Cu^{2+} + 2e^-$, so $n(Cu) = 0.2238 \div 2 = 0.1119$ mol

5. $m = nM = 0.1119 \times 63.55 = 7.11$ g. The anode loses 7.11 g.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 1-2 of 7

1. Given a cathode that gains 0.850 g of silver in 15.0 minutes from $Ag^+ + e^- \rightarrow Ag$ (M = 107.87 g/mol). Find the constant current.

n(Ag) = 0.850 ÷ 107.87 = 0.00788 mol. n(e^-) = 0.00788 mol. q = 760.4 C. t = 900 s. I = 760.4 ÷ 900 = 0.845 A.

2. Given 1.50 A for 30.0 minutes and 0.485 g of a metal deposited (M = 52.00 g/mol). Find the charge on the metal ion.

q = 1.50 × 1800 = 2700 C. n(e^-) = 0.02798 mol. n(metal) = 0.485 ÷ 52.00 = 0.00933 mol. z = 0.02798 ÷ 0.00933 = 3.00.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 3-4 of 7

3. Given a zinc and copper voltaic cell that delivers 0.800 A for 45.0 minutes (M(Zn) = 65.41; M(Cu) = 63.55 g/mol). Find the mass of zinc lost and the mass of copper gained.

q = 0.800 × 2700 = 2160 C. n(e^-) = 0.02238 mol. n(Zn) = n(Cu) = 0.01119 mol. Zinc lost: 0.732 g. Copper gained: 0.711 g.

4. Given $Na_2SO_4(aq)$ with inert electrodes and 3.00 A for 20.0 minutes. Find the moles of hydrogen and oxygen formed.

q = 3.00 × 1200 = 3600 C. n(e^-) = 0.037306 mol. n(H_2) = 0.0187 mol and n(O_2) = 0.00933 mol.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 5-6 of 7

5. Given a cathode of 20.00 g that has a mass of 21.20 g after copper plating at 2.00 A. Find the time in minutes (M(Cu) = 63.55 g/mol).

Use Method: how to solve a Faraday problem when the unknown changes. Mass deposited = 1.20 g. n(Cu) = 0.01888 mol. n(e^-) = 0.03777 mol. q = 3644 C. t = 1822 s = 30.4 min.

6. Given a copper anode in a refining cell with 5.00 A for 90.0 minutes. Find the mass lost by the anode (M(Cu) = 63.55 g/mol).

q = 5.00 × 5400 = 2.70 × 10^4 C. n(e^-) = 0.2798 mol. n(Cu) = 0.1399 mol. m = 8.89 g.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 7-7 of 7

7. Challenge: Given 2.00 A for 25.0 minutes, which deposits 0.912 g of an unknown metal from a solution of its 2+ ions. Find the molar mass of the metal in g/mol. Then name the metal, using the periodic table in the Data Booklet. Use F = 9.65 × 10^4 C/mol e^-.

The unknown is the molar mass, so run the chain from the charge to moles of metal. Time: t = 25.0 × 60 = 1500 s. Charge: q = It = 2.00 × 1500 = 3000 C. Moles of electrons: n(e^-) = 3000 ÷ (9.65 × 10^4) = 0.031088 mol. The ion is 2+, so n(metal) = 0.031088 ÷ 2 = 0.015544 mol. Molar mass: M = m ÷ n = 0.912 g ÷ 0.015544 mol = 58.7 g/mol. The periodic table gives nickel, 58.69 g/mol.