B10 · Faraday's law

Completed notes

Goals: Find the charge from current and time, with the time in seconds. Use the Faraday constant and a half-reaction to find moles of metal. Find the mass deposited, or the time needed, in an electrolysis.

Slide 1 of 8 · Current is charge per second

Electric charge is measured in **coulombs**, C. Current, I, is the charge that passes a point every **second**.

Current is measured in **amperes**, A. One ampere means **one coulomb** each second.

Charge equals current times time: q = **It**. Here q is the charge in coulombs, I is the current in amperes and t is the time in **seconds**.

If the time is in minutes or hours, change it to seconds first. 1 min = **60** s and 1 h = **3600** s.

Example. A current of 4.00 A runs through an electrolytic cell for 5.00 minutes. How much charge moves, in coulombs?

1. Time in seconds: 5.00 × 60 = 300 s

2. q = It = 4.00 A × 300 s

3. q = 1200 C

Try it. A current of 2.50 A runs through an electrolytic cell for 12.0 minutes. How much charge moves, in coulombs?

1. Time in seconds: 12.0 × 60 = 720 s

2. q = It = 2.50 A × 720 s

3. q = 1.80×10^3 C

Slide 2 of 8 · Charge to moles of electrons

One mole of electrons carries **9.65 × 10⁴ C**. This is the **Faraday constant**, F, printed in the Data Booklet on p. 3.

F = 9.65 × 10⁴ C/mol e^-. The unit says coulombs for **each mole** of electrons.

To change coulombs to moles of electrons, **divide** by F: $n(e^-) = \mathbf{q \div F}$. Here $n(e^-)$ is the moles of electrons and q is the charge.

Coulombs divided by coulombs per mole leaves **moles**. A small charge should give a **small** number of moles.

Example. A charge of 1500 C moves through a cell. How many moles of electrons is that? Use F = 9.65 × 10⁴ C/mol e^-.

1. $n(e^-) = q \div F = 1500\ C \div (9.65 \times 10^4)$

2. $n(e^-) = 0.0155$ mol

3. Check: 1500 C is a small piece of 9.65 × 10⁴ C, so a small answer makes sense.

Try it. A charge of 4.80 × 10³ C moves through a cell. How many moles of electrons is that? Use F = 9.65 × 10⁴ C/mol e^-.

1. $n(e^-) = q \div F = 4800\ C \div (9.65 \times 10^4)$

2. $n(e^-) = 0.0497$ mol

Slide 3 of 8 · Electrons to moles of metal

The coefficient of e^- in the reduction half-reaction gives the mole ratio. In $Cu^{2+} + 2e^- \rightarrow Cu$, **2** mol of electrons make **1** mol of copper.

So $n(Cu) = n(e^-) \div 2$. Here n(Cu) is the moles of copper.

For silver, $Ag^+ + e^- \rightarrow Ag$, the ratio is **1 to 1**. For aluminium, $Al^{3+} + 3e^- \rightarrow Al$, divide by **3**.

The ratio is set by the **charge on the ion**. Write the half-reaction every time and never guess the ratio.

Example. A cell supplies 0.0155 mol of electrons to the cathode, where $Cu^{2+} + 2e^- \rightarrow Cu$. How many moles of copper form?

1. The half-reaction shows 2 mol e^- make 1 mol Cu.

2. $n(Cu) = n(e^-) \div 2 = 0.0155 \text{ mol} \div 2$

3. $n(Cu) = 0.00775 \text{ mol}$

Try it. A cell supplies 0.0450 mol of electrons to the cathode, where $Al^{3+} + 3e^- \rightarrow Al$. How many moles of aluminium form?

1. The half-reaction shows 3 mol e^- make 1 mol Al.

2. $n(Al) = n(e^-) \div 3 = 0.0450 \text{ mol} \div 3$

3. $n(Al) = 0.0150 \text{ mol}$

Slide 4 of 8 · The full chain to a mass

The Faraday chain is: current and time give **charge**, charge gives **moles of electrons**, the half-reaction gives moles of metal, and molar mass gives mass.

First change the time to seconds and find q = It. Then n(e^-) = **q ÷ F**. Then divide by the **coefficient of e^-** to find n(metal).

Finally m = **nM**. Here n is the moles of metal and M is the molar mass in **g/mol**, from the periodic table.

The chain as one formula is m = (I × t × M) ÷ (z × F). Here z is the **electrons per ion** in the half-reaction, and t is in **seconds**.

Example. Zinc is plated from Zn^{2+} solution by a current of 2.50 A for 25.0 minutes. What mass of zinc forms at the cathode? M(Zn) = 65.41 g/mol.

1. Time in seconds: 25.0 × 60 = 1500 s

2. q = It = 2.50 × 1500 = 3750 C

3. n(e^-) = 3750 ÷ (9.65 × 10^4) = 0.038860 mol

4. Zn^{2+} + $2e^-$ → Zn, so n(Zn) = 0.038860 ÷ 2 = 0.019430 mol

5. m = nM = 0.019430 × 65.41 = 1.27 g

Try it. Copper is plated from Cu^{2+} solution by a current of 3.00 A for 30.0 minutes. What mass of copper forms at the cathode? M(Cu) = 63.55 g/mol.

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

2. q = It = 3.00 × 1800 = 5400 C

3. n(e^-) = 5400 ÷ (9.65 × 10^4) = 0.05596 mol

4. Cu^{2+} + $2e^-$ → Cu, so n(Cu) = 0.05596 ÷ 2 = 0.02798 mol

5. m = nM = 0.02798 × 63.55 = 1.78 g

Slide 5 of 8 · Method: how to find a mass from current and time

Step 1. Change the time to **seconds**. Multiply minutes by 60.

Step 2. Find the charge: q = **It**.

Step 3. Find the moles of electrons: $n(e^-) = q \div \mathbf{F}$, with F = 9.65×10^4 C/mol e^-.

Step 4. Use the **half-reaction** ratio to find the moles of the substance.

Step 5. Find the mass: m = **nM**.

Step 6. Round **once**, at the end, to three significant digits.

Example. Use the method on 2.50 A for 25.0 minutes plating zinc from Zn^{2+} solution. Find the mass of zinc.

1. Step 1: $25.0 \times 60 = 1500$ s.

2. Step 2: $q = 2.50 \times 1500 = 3750$ C.

3. Step 3: $n(e^-) = 3750 \div (9.65 \times 10^4) = 0.038860$ mol.

4. Step 4: $n(Zn) = 0.038860 \div 2 = 0.019430$ mol.

5. Steps 5 and 6: $m = 0.019430 \times 65.41 = 1.2709$ g, which rounds to 1.27 g.

Try it. Use the method on 1.50 A for 20.0 minutes plating silver from Ag^+ solution. M(Ag) = 107.87 g/mol. Find the mass of silver.

1. $q = 1.50 \times 1200 = 1800$ C.

2. $n(e^-) = 1800 \div (9.65 \times 10^4) = 0.018653$ mol. $n(Ag) = 0.018653$ mol, because $Ag^+ + e^- \rightarrow Ag$.

3. $m = 0.018653 \times 107.87 = 2.0121$ g, which rounds to 2.01 g.

Slide 6 of 8 · Your turn: silver

Hours must be changed to seconds in two steps: 1.00 h = **60** min = **3600** s.

Charge: q = It = 0.500 A × 3600 s = **1800 C**. Then n(e⁻) = q ÷ F = **0.01865** mol.

In $Ag^+ + e^- \rightarrow Ag$ the ratio is **1 to 1**, so n(Ag) = n(e⁻). Show this step even when the ratio is 1 to 1.

Finally m = n × M = 0.01865 mol × **107.87 g/mol**. Check that the units give grams: mol × g/mol leaves **grams**.

Example. A silver nitrate solution carries 0.500 A for 1.00 hour. The half-reaction is $Ag^+ + e^- \rightarrow Ag$. What mass of silver forms at the cathode? M(Ag) = 107.87 g/mol.

1. Time in seconds: 1.00 × 3600 = 3600 s

2. q = It = 0.500 × 3600 = 1800 C

3. $n(e^-) = 1800 \div (9.65 \times 10^4) = 0.01865$ mol

4. Ratio 1 to 1, so n(Ag) = 0.01865 mol

5. m = nM = 0.01865 × 107.87 = 2.01 g

Try it. A copper(II) sulfate solution carries 0.750 A for 45.0 minutes. The half-reaction is $Cu^{2+} + 2e^- \rightarrow Cu$. What mass of copper forms at the cathode? M(Cu) = 63.55 g/mol.

1. Time in seconds: 45.0 × 60 = 2700 s

2. q = It = 0.750 × 2700 = 2025 C

3. $n(e^-) = 2025 \div (9.65 \times 10^4) = 0.02098$ mol

4. n(Cu) = 0.02098 ÷ 2 = 0.01049 mol

5. m = nM = 0.01049 × 63.55 = 0.667 g

Slide 7 of 8 · Running the chain backward

To run the chain backward, start at the given **mass** and find moles with n = **m ÷ M**. Here m is the mass in grams and M is the molar mass.

Go from metal to electrons by **multiplying** by the coefficient of e^-. For $Al^{3+} + 3e^- \rightarrow Al$, $n(e^-) = \mathbf{3} \times n(Al)$.

Change moles of electrons to charge by multiplying by **F**. Then find the time with t = **q ÷ I**.

Coulombs divided by amperes leaves **seconds**. Divide by **60** for minutes if the question asks for minutes.

Example. We want 1.00 g of aluminium from Al^{3+} using a steady current of 5.00 A. How long must the current run, in seconds and in minutes? M(Al) = 26.98 g/mol.

1. n(Al) = 1.00 ÷ 26.98 = 0.03706 mol

2. $Al^{3+} + 3e^- \rightarrow Al$, so $n(e^-) = 3 \times 0.03706 = 0.1112$ mol

3. $q = 0.1112 \times (9.65 \times 10^4) = 1.073 \times 10^4$ C

4. t = q ÷ I = $1.073 \times 10^4 \div 5.00 = 2146$ s

5. t = 2146 ÷ 60 = 35.8 min

Try it. We want 1.50 g of zinc from Zn^{2+} using a steady current of 3.00 A. How long must the current run, in minutes? M(Zn) = 65.41 g/mol.

1. n(Zn) = 1.50 ÷ 65.41 = 0.02293 mol

2. $Zn^{2+} + 2e^- \rightarrow Zn$, so $n(e^-) = 2 \times 0.02293 = 0.04586$ mol

3. $q = 0.04586 \times (9.65 \times 10^4) = 4425$ C

4. t = q ÷ I = 4425 ÷ 3.00 = 1475 s

5. t = 1475 ÷ 60 = 24.6 min

Slide 8 of 8 · What the calculation assumes

The calculation assumes **all** the charge makes the wanted metal. In real cells some charge goes to **side reactions**, so a little less metal forms.

Mass deposited is the mass the cathode **gained**: the final mass **minus** the starting mass.

The same chain also works in a **voltaic** cell. The electrons in the wire come from the **anode** half-reaction.

In a zinc and copper cell, $Zn \rightarrow Zn^{2+} + \mathbf{2e^-}$ and $Cu^{2+} + 2e^- \rightarrow \mathbf{Cu}$. The same electrons do both jobs.

Example. A cathode has a mass of 31.40 g before plating and 32.15 g after plating. What is the mass deposited? What does the calculation assume about the charge?

1. Mass deposited = final mass − starting mass

2. 32.15 g − 31.40 g = 0.75 g

3. The calculation assumes all the charge made the wanted metal, with no side reactions.

Try it. In a zinc and copper voltaic cell, 0.0200 mol of electrons pass through the wire. How many moles of zinc are oxidized at the anode, and how many moles of copper deposit on the cathode?

1. $Zn \rightarrow Zn^{2+} + 2e^-$, so n(Zn) = 0.0200 ÷ 2 = 0.0100 mol

2. $Cu^{2+} + 2e^- \rightarrow Cu$, so n(Cu) = 0.0200 ÷ 2 = 0.0100 mol

3. The same electrons do both jobs, so the moles are equal.

B10 · Worked practice

Faraday's law | Questions 1-2 of 7

1. Given a current of 4.00 A for 30.0 minutes. Find the charge that moves, in coulombs.

The time is 30.0 × 60 = 1800 s. q = It = 4.00 A × 1800 s = 7.20×10^3 C.

2. Given a charge of 7.20×10^3 C. Find the moles of electrons, using F = 9.65×10^4 C/mol e^-.

$n(e^-) = q \div F = 7.20 \times 10^3 \div (9.65 \times 10^4) = 0.0746$ mol.

B10 · Worked practice

Faraday's law | Questions 3-4 of 7

3. Given 0.0746 mol of electrons and the half-reaction $Cu^{2+} + 2e^- \rightarrow Cu$. Find the moles of copper formed and the mass (M = 63.55 g/mol).

n(Cu) = 0.0746 ÷ 2 = 0.0373 mol. m = 0.0373 × 63.55 = 2.37 g.

4. Given 2.00 A for 45.0 minutes and the half-reaction $Zn^{2+} + 2e^- \rightarrow Zn$ (M = 65.41 g/mol). Find the mass of zinc deposited.

Use Method: how to find a mass from current and time. q = 2.00 × 2700 = 5400 C. $n(e^-) = 5400 \div (9.65 \times 10^4) = 0.0560$ mol. n(Zn) = 0.0280 mol. m = 0.0280 × 65.41 = 1.83 g.

B10 · Worked practice

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5. Given 1.50 A for 2.00 hours and $Ag^+ + e^- \rightarrow Ag$ (M = 107.87 g/mol). Find the mass of silver deposited.

$q = 1.50 \times 7200 = 1.08 \times 10^4$ C. $n(e^-) = 0.112$ mol. $n(Ag) = 0.112$ mol. $m = 0.112 \times 107.87 = 12.1$ g.

6. Given that 2.00 g of copper must form from Cu^{2+} at a steady 2.50 A (M = 63.55 g/mol). Find the time needed in minutes.

$n(Cu) = 2.00 \div 63.55 = 0.03147$ mol. $n(e^-) = 0.06294$ mol. $q = 0.06294 \times (9.65 \times 10^4) = 6074$ C. $t = 6074 \div 2.50 = 2430$ s = 40.5 min.

B10 · Worked practice

Faraday's law | Questions 7-7 of 7

7. Challenge: Given two electrolytic cells joined in a row, so the same current passes through both. One cell plates silver ($Ag^+ + e^- \rightarrow Ag$, M = 107.87 g/mol). The other plates copper ($Cu^{2+} + 2e^- \rightarrow Cu$, M = 63.55 g/mol). The current is 1.25 A for 20.0 minutes. Find the mass of each metal deposited, in grams. Use $F = 9.65 \times 10^4$ C/mol e^-.

Time: $t = 20.0 \times 60 = 1200$ s. Charge: $q = It = 1.25 \times 1200 = 1500$ C. The same charge passes through both cells. Moles of electrons: $n(e^-) = 1500 \div (9.65 \times 10^4) = 0.015544$ mol. Silver, ratio 1 to 1: $n(Ag) = 0.015544$ mol, so $m = 0.015544 \times 107.87 = 1.68$ g. Copper, ratio 2 e^- to 1 Cu: $n(Cu) = 0.007772$ mol, so $m = 0.007772 \times 63.55 = 0.494$ g. Equal charge does not give equal mass.