B10 · Faraday's law

Completed notes | Notes 1/3

Find the charge from current and time, change it to moles of electrons, use a half-reaction to find moles of metal, and find a mass or a time.

1. Charge and moles of electrons

Current is charge per second

Electric charge is the amount of electricity that moves. We measure it in **coulombs**, C. Current is how much charge passes a point each second.

Current is measured in amperes, A. One ampere is one **coulomb** each second. So charge equals current times time: q = **It**.

An ampere already has seconds built in, so the time must be in **seconds**. For 2.50 A for 10.0 min, the time is 10.0 × 60 = **600** s.

Then q = 2.50 × 600 = **1500 C**. On the diploma: leaving the time in minutes or hours is one of the most common wrong options.

Watch: Current is coulombs per second (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 3:32)

Watch: Change hours and minutes to seconds (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 4:16)

The Faraday constant

Chemistry counts in moles, so we need a number that links coulombs to moles. One mole of electrons carries **9.65×10^4** C.

This is the Faraday constant, F. Look it up in the **data booklet**. Think of it as a molar mass for charge.

To change coulombs into moles of electrons, divide: $n(e^-) = q \div \mathbf{F}$. The coulombs cancel and moles are left.

For 1500 C, $n(e^-) = 1500 \div (9.65 \times 10^4) =$ **0.0155 mol**. That is a small piece of one mole, so a small answer makes sense.

On the diploma: a common wrong option multiplies by F when it should divide.

Watch: Faraday’s constant: the charge on one mole of electrons (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 4:59)

Watch: Moles of electrons = current × time ÷ F (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 6:22)

Given 2.50 A for 10.0 min. Time in seconds: 600 s; q = 1500 C

$n(e^-) = q \div F = 1500 \div (9.65 \times 10^4) = 0.0155$ mol

$Cu^{2+} + 2e^- \rightarrow Cu$, so $n(Cu) = 0.00777$ mol

$m = n \times M = 0.00777 \times 63.55 = 0.494$ g

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2. Moles of metal and mass

Use the half-reaction as the ratio

Moles of electrons are not yet moles of metal. The reduction half-reaction at the cathode gives the ratio, because it shows how many electrons one atom needs.

In $Cu^{2+} + 2e^- \rightarrow Cu$, **2** mol of electrons make 1 mol of copper. In $Ag^+ + e^- \rightarrow Ag$, 1 mol of electrons makes **1** mol of silver.

In $Al^{3+} + 3e^- \rightarrow Al$, three electrons make one atom, so divide the moles of electrons by **3**. The ion charge sets the ratio.

Always write the half-reaction first. On the diploma: a wrong option often uses the wrong ratio, like 1 to 1 for copper.

Watch: Use the half-reaction, not the whole equation (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 1:24)

Watch: The coefficients give electrons to metal (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 12:31)

The full chain

Let’s put the links together. Zinc is plated by 3.00 A for 25.0 min. The time is 25.0 × 60 = **1500** s, so q = 3.00 × 1500 = **4500** C.

Then $n(e^-) = 4500 \div (9.65 \times 10^4) =$ **0.0466** mol. In $Zn^{2+} + 2e^- \rightarrow Zn$, each zinc atom needs two electrons, so n(Zn) = **0.0233** mol.

Mass is moles times molar mass: m = 0.0233 × 65.41 = **1.52 g**. The path is charge, moles of electrons, moles of metal, then mass.

As one formula, m = **ItM** ÷ (zF). M is the molar mass and z is the number of electrons for one **ion**.

On the diploma: if a formula is given, check that t is in seconds and that z matches the half-reaction.

Watch: Example: zinc plated on a nail (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 7:45)

Watch: Change minutes to seconds (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 10:27)

Watch: Moles to mass (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 13:51)

Watch: Second example: copper on nickel (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 15:55)

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3. Running the chain backward

Find a time

Now run the chain backward. We want 1.00 g of aluminium at 5.00 A. Start at the given, the mass, and find the moles first.

n(Al) = 1.00 ÷ 26.98 = **0.03706** mol. In $Al^{3+} + 3e^- \rightarrow Al$, each atom needs three electrons. Going backward we multiply: $n(e^-)$ = **0.1112** mol.

Moles of electrons to coulombs means multiply by F: $q = 0.1112 \times (9.65 \times 10^4)$ = **1.073×10^4** C.

Since q = It, the time is $t = q \div I = 1.073 \times 10^4 \div 5.00$ = **2146** s. Divide by 60 to get **35.8** min.

On the diploma: check which unit the question wants for time. A wrong option divides by three where you should multiply.

What the calculation assumes

The calculation assumes all the **charge** makes the product in the half-reaction. Real cells can lose some charge to **side** reactions, which are other unwanted reactions.

Mass deposited is the **gain** in mass of the cathode, not its **final** mass. The cathode already had some mass before the current started.

On the diploma: if a question gives two weighings of the cathode, use the difference, not the final reading.

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Faraday's law | Questions 1-2 of 12

State symbols: (s) solid, (aq) dissolved in water, (l) liquid, (g) gas. Use the supplied reaction and conditions. B4 values name the Alberta Chemistry 30 Data Booklet table where used.

B10-P01 · Easy · A current of 3.50 A flows for 12.0 minutes. Find the charge that moves.

1. Change the time to seconds: 12.0 × 60 = 720 s.

2. q = It = 3.50 A × 720 s = 2.52 × 10^3 C.

B10-P02 · Easy · A student calculates the charge for 2.00 A and 15 minutes as 30 C. Find the correct charge and explain the error.

1. The student used 15 as the time in seconds. The ampere needs the time in seconds.

2. The correct time is 900 s, so q = 2.00 × 900 = 1.80 × 10^3 C.

B10 · Worked practice

Faraday's law | Questions 3-4 of 12

B10-P03 · Easy · A charge of 2.52×10^3 C passes through a cell. Find the moles of electrons ($F = 9.65 \times 10^4$ C/mol e^-).

1. Divide the charge by the Faraday constant.

2. $n(e^-) = 2.52 \times 10^3 \div 9.65 \times 10^4 = 0.0261$ mol.

B10-P04 · Easy · Explain in one or two sentences what the Faraday constant is and why it works like a molar mass.

1. F is the charge carried by one mole of electrons, 9.65×10^4 C/mol.

2. A molar mass changes moles to grams. F changes moles of electrons to coulombs, and the other way.

B10 · Worked practice

Faraday's law | Questions 5-6 of 12

B10-P05 · Medium · Given 0.0261 mol of electrons and the half-reaction $Cu^{2+} + 2e^- \rightarrow Cu$. Find the moles of copper and its mass (M = 63.55 g/mol).

1. Two moles of electrons make one mole of copper, so n(Cu) = 0.0261 ÷ 2 = 0.01306 mol.

2. m = 0.01306 × 63.55 = 0.830 g.

B10-P06 · Medium · Given 0.0261 mol of electrons. Find the moles of aluminium formed from $Al^{3+} + 3e^- \rightarrow Al$, and compare with silver from $Ag^+ + e^- \rightarrow Ag$.

1. Aluminium: n(Al) = 0.0261 ÷ 3 = 0.00870 mol.

2. Silver: n(Ag) = 0.0261 ÷ 1 = 0.0261 mol. The same electrons make three times as many silver atoms as aluminium atoms.

B10 · Worked practice

Faraday's law | Questions 7-8 of 12

B10-P07 · Medium · Given 1.80 A for 40.0 minutes and $Ag^{+} + e^{-} \rightarrow Ag$ (M = 107.87 g/mol). Find the mass of silver deposited.

1. q = 1.80 × 2400 = 4320 C, and $n(e^{-}) = 4320 \div (9.65 \times 10^{4}) = 0.04477$ mol.

2. The ratio is 1 to 1, so n(Ag) = 0.04477 mol and m = 0.04477 × 107.87 = 4.83 g.

B10-P08 · Medium · Given 2.40 A for 50.0 minutes and $Zn^{2+} + 2e^{-} \rightarrow Zn$ (M = 65.41 g/mol). Find the mass of zinc deposited.

1. q = 2.40 × 3000 = 7200 C, and $n(e^{-}) = 7200 \div (9.65 \times 10^{4}) = 0.07461$ mol.

2. n(Zn) = 0.07461 ÷ 2 = 0.03731 mol and m = 0.03731 × 65.41 = 2.44 g.

B10 · Worked practice

Faraday's law | Questions 9-10 of 12

B10-P09 · Challenging · Given $Ni^{2+} + 2e^- \rightarrow Ni$. Use $m = ItM \div (zF)$ to find the mass of nickel from 1.20 A for 50.0 minutes (M = 58.69 g/mol).

1. The time is 3000 s and z = 2.

2. $m = (1.20 \times 3000 \times 58.69) \div (2 \times (9.65 \times 10^4)) = 1.09$ g.

B10-P10 · Challenging · Given that 2.00 g of copper must form from Cu^{2+} at a steady 4.00 A (M = 63.55 g/mol). Find the time needed in minutes.

1. $n(Cu) = 2.00 \div 63.55 = 0.03147$ mol, and $n(e^-) = 2 \times 0.03147 = 0.06294$ mol.

2. $q = 0.06294 \times (9.65 \times 10^4) = 6.07 \times 10^3$ C. Then $t = 6074 \div 4.00 = 1519$ s = 25.3 min.

B10 · Worked practice

Faraday's law | Questions 11-12 of 12

B10-P11 · Challenging · Given that 1.00 g of aluminium must form from $Al^{3+} + 3e^- \rightarrow Al$ at 2.50 A. Find the time in minutes (M = 26.98 g/mol).

1. n(Al) = 1.00 ÷ 26.98 = 0.03706 mol, and $n(e^-) = 3 \times 0.03706 = 0.1112$ mol.

2. $q = 0.1112 \times (9.65 \times 10^4) = 1.073 \times 10^4$ C. Then t = 10 730 ÷ 2.50 = 4292 s = 71.5 min.

B10-P12 · Challenging · A cathode has a mass of 18.40 g before plating and 19.10 g after. Find the mass deposited, and say why we use it and not 19.10 g.

1. The mass deposited is the gain: 19.10 − 18.40 = 0.70 g.

2. Only the gain was made by the current. The starting mass was already on the cathode. Real cells can also lose some charge to side reactions.