B10 · Faraday's law

Student notes - complete the underlined spaces | 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.

Name: _____ Date: _____

1. Charge and moles of electrons

Current is charge per second

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

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

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

Then q = 2.50 × 600 = _____. 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 _____ C.

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

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

For 1500 C, $n(e^-) = 1500 \div (9.65 \times 10^4)$ = _____. 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: _____; q = _____

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

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

$m = n \times M = 0.00777 \times 63.55$ = _____

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

In $Al^{3+} + 3e^- \rightarrow Al$, three electrons make one atom, so divide the moles of electrons by _____. 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 = _____ s, so q = 3.00 × 1500 = _____ C.

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

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

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

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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Student notes - complete the underlined spaces | Notes 3/3

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

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

Since q = It, the time is $t = q \div I = 1.073 \times 10^4 \div 5.00$ = _____ s. Divide by 60 to get _____ 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 _____ makes the product in the half-reaction. Real cells can lose some charge to _____ reactions, which are other unwanted reactions.

Mass deposited is the _____ in mass of the cathode, not its _____ 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.

B10 · Practice

Faraday's law | Questions 1-3 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.

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.

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^-).

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

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).

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 $Ag^+ + e^- \rightarrow Ag$.

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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.

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.

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).

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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.

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).

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.