B11 · Faraday's law with the unknown chosen

Completed notes | Notes 1/3

Start the Faraday chain from the given quantities to find a current, a time, an ion charge or a mass, and use the same charge in cells in series.

1. One chain, any unknown

Choose the direction

In B10 we built one chain: current and time give **charge**, charge gives moles of electrons, the half-reaction gives moles of **metal**, and the molar mass gives the **mass**.

Every Faraday problem uses this same chain. Only the given end and the asked end change.

So write the **givens** and the question first. Then start at a given and walk toward the **unknown**.

If the mass is given, go **backward**. Multiply by the coefficient of the electrons and then by **F**. Going forward, you divide, because each step reverses.

On the diploma: many wrong options divide when they should multiply, or the other way round.

Watch: Two kinds of question (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 2:04)

Watch: Working backward from the mass (Alberta Chemistry Teacher, “Chemistry 30 - Cell stoichiometry theory and examples”, 20:02)

Find a current

A copper cathode gains 1.25 g in 40.0 min from Cu^{2+}. The mass is given, so start there: n(Cu) = 1.25 ÷ 63.55 = **0.01967** mol.

$Cu^{2+} + 2e^- \rightarrow Cu$, so multiply by 2: $n(e^-)$ = **0.03934** mol. Moles to coulombs means multiply by F.

q = 0.03934 × (9.65 × 10^4) = **3796** C. The time is 40.0 × 60 = 2400 s, so I = q ÷ t = 3796 ÷ 2400 = **1.58 A**.

Leaving the time in minutes would make the current 60 times too big. On the diploma: write givens with units, like t = 40.0 min, so you see what to convert.

Zn/Cu cell, 0.500 A for 30.0 min. q = 900 C

$n(e^-) = 900 \div (9.65 \times 10^4) = 0.00933$ mol

$Zn \rightarrow Zn^{2+} + 2e^-$, so zinc lost = 0.305 g

$Cu^{2+} + 2e^- \rightarrow Cu$, so copper gained = 0.296 g

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Completed notes | Notes 2/3

2. Ion charge and the voltaic cell

Find the ion charge

Sometimes the ion is the mystery. Let z be the number of electrons each ion gains. Given 2.00 A for 20.0 min and 0.790 g of a metal with M = 63.55 g/mol.

The time is 1200 s, so q = **2400** C and $n(e^-) = 2400 \div (9.65 \times 10^4)$ = **0.02487** mol.

n(metal) = 0.790 ÷ 63.55 = **0.01243** mol. Dividing electrons by atoms gives electrons per atom: z = **2.00**, so the ion is M^{2+}.

z must be a **whole** number, because ion charges are whole numbers. Round to the nearest integer. On the diploma: do not report 2.0007.

The chain in a voltaic cell

An electrode is a piece of metal where a half-reaction happens. In a voltaic cell, the electrons in the wire come from the **anode** and go to the **cathode**.

So the same chain links the charge to the mass at both electrodes. A Zn/Cu cell gives 0.500 A for 30.0 min: q = 0.500 × 1800 = **900** C.

Then $n(e^-) = 900 \div (9.65 \times 10^4)$ = **0.009326** mol. Both half-reactions use 2 electrons per atom, so n(Zn) = n(Cu) = **0.004663** mol.

The zinc anode loses 0.004663 × 65.41 = **0.305** g. The copper cathode gains 0.004663 × 63.55 = **0.296** g. The moles are equal, but the molar **masses** are different.

On the diploma: say which electrode loses mass and which gains mass.

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Completed notes | Notes 3/3

3. Gases, mass gained and the anode

Gases from water

$Na_2SO_4(aq)$ is electrolyzed with inert electrodes, which carry electrons but do not react. At 2.00 A for 30.0 min, q = **3600** C and $n(e^-)$ = **0.0373** mol.

Reduction: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$. It makes one H_2 for two electrons, so $n(H_2)$ = 0.0373 ÷ 2 = **0.01865** mol.

Oxidation: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$. It makes one O_2 for four electrons, so $n(O_2)$ = 0.0373 ÷ 4 = **0.00933** mol.

The same electrons pass through the whole circuit. Hydrogen forms in a ratio of **2** to **1** with oxygen, as in $2H_2O \rightarrow 2H_2 + O_2$.

On the diploma: a wrong option divides by 2 for oxygen. Use the 4 from the oxidation half-reaction.

Mass gained and mass lost

Use the mass **gained** by the cathode. A cathode that goes from 14.20 g to 14.85 g gained **0.65** g.

With 1.50 A, n(Cu) = 0.65 ÷ 63.55 = 0.01023 mol and $n(e^-)$ = 0.02046 mol. Then q = 1974 C and t = 1974 ÷ 1.50 = **1316** s, or 21.9 min.

In copper refining, the anode loses mass: $Cu \rightarrow Cu^{2+} + 2e^-$. For 10.0 A for 2.00 h, the time is 7200 s and q = **7.20×10^4** C.

Then $n(e^-)$ = 0.746 mol and n(Cu) = 0.373 mol, so the anode loses **23.7** g of copper. If no electrons go to side reactions, the pure cathode gains the same amount.

B11 · Worked practice

Faraday's law with the unknown chosen | 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.

B11-P01 · Easy · A cathode gains 0.850 g of silver in 15.0 minutes from $Ag^+ + e^- \rightarrow Ag$ (M = 107.87 g/mol). Find the constant current.

1. n(Ag) = 0.850 ÷ 107.87 = 0.00788 mol, so $n(e^-)$ = 0.00788 mol and q = 0.00788 × (9.65×10^4) = 760 C.

2. The time is 900 s, so I = 760 ÷ 900 = 0.845 A.

B11-P02 · Easy · A cathode gains 2.50 g of nickel in 30.0 minutes from $Ni^{2+} + 2e^- \rightarrow Ni$ (M = 58.69 g/mol). Find the current.

1. n(Ni) = 2.50 ÷ 58.69 = 0.04259 mol, and $n(e^-)$ = 0.08519 mol. Then q = 0.08519 × (9.65×10^4) = 8221 C.

2. The time is 1800 s, so I = 8221 ÷ 1800 = 4.57 A.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 3-4 of 12

B11-P03 · Easy · A student finds the current by dividing the charge by the time in minutes. Explain the error and the fix.

1. The ampere is coulombs per second, so the time must be in seconds.

2. Divide the charge by the time in seconds. Otherwise the current comes out 60 times too large.

B11-P04 · Easy · 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.

1. q = 1.50 × 1800 = 2700 C, and n(e^-) = 2700 ÷ (9.65×10^4) = 0.02798 mol.

2. n(metal) = 0.485 ÷ 52.00 = 0.00933 mol. Then z = 0.02798 ÷ 0.00933 = 3.00, so the ion is M^{3+}.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 5-6 of 12

B11-P05 · Medium · A calculation of z gives 2.96. Explain why we state z = 3 and what the small difference tells us.

1. z is the number of electrons for each ion, so it must be a whole number.

2. The small difference comes from rounding and measurement error, so we round to the nearest whole number.

B11-P06 · Medium · A zinc and copper voltaic cell gives 0.800 A for 45.0 minutes. Find the mass of zinc lost from the anode (M = 65.41 g/mol).

1. q = 0.800 × 2700 = 2160 C, and $n(e^-) = 2160 \div (9.65 \times 10^4) = 0.02238$ mol.

2. $Zn \rightarrow Zn^{2+} + 2e^-$, so n(Zn) = 0.01119 mol and m = 0.01119 × 65.41 = 0.732 g.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 7-8 of 12

B11-P07 · Medium · For the same cell, find the mass of copper gained at the cathode (M = 63.55 g/mol), and explain why the two masses are different.

1. $Cu^{2+} + 2e^- \rightarrow Cu$, so n(Cu) = 0.01119 mol and m = 0.01119 × 63.55 = 0.711 g.

2. The moles are equal because each metal uses two electrons for each atom. The molar masses are different, so the masses are different.

B11-P08 · Medium · $Na_2SO_4(aq)$ is electrolyzed with inert electrodes at 3.00 A for 20.0 minutes. Find the moles of hydrogen and oxygen formed.

1. q = 3.00 × 1200 = 3600 C, and $n(e^-) = 3600 \div (9.65 \times 10^4) = 0.0373$ mol.

2. $n(H_2) = 0.0373 \div 2 = 0.01865$ mol and $n(O_2) = 0.0373 \div 4 = 0.00933$ mol.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 9-10 of 12

B11-P09 · Challenging · For the same electrolysis, find the mass of hydrogen formed (M = 2.02 g/mol) and state the ratio of hydrogen to oxygen by moles.

1. $m(H_2) = 0.01865 \times 2.02 = 0.0377$ g.

2. The ratio is 0.01865 : 0.00933, which is 2 : 1, and it agrees with $2H_2O \rightarrow 2H_2 + O_2$.

B11-P10 · Challenging · A cathode of 20.00 g has a mass of 21.20 g after copper plating at 2.00 A. Find the time in minutes (M = 63.55 g/mol).

1. The mass deposited is 21.20 − 20.00 = 1.20 g, so n(Cu) = 0.018883 mol and $n(e^-)$ = 0.037766 mol.

2. $q = 0.037766 \times (9.65 \times 10^4) = 3644$ C. Then t = 3644 ÷ 2.00 = 1822 s = 30.4 min.

B11 · Worked practice

Faraday's law with the unknown chosen | Questions 11-12 of 12

B11-P11 · Challenging · A copper anode in a refining cell runs at 5.00 A for 90.0 minutes. Find the mass lost by the anode (M = 63.55 g/mol).

1. $q = 5.00 \times 5400 = 2.70 \times 10^4$ C, and $n(e^-) = 0.2798$ mol.

2. $Cu \rightarrow Cu^{2+} + 2e^-$, so $n(Cu) = 0.1399$ mol and $m = 0.1399 \times 63.55 = 8.89$ g.

B11-P12 · Challenging · A student uses the final mass of a cathode, 21.20 g, to find a plating time. Explain the error and say whether the time is too large or too small.

1. Only the gain in mass was made by the current. The starting mass was already on the cathode.

2. Using the final mass makes the moles, the charge and the time too large.