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

B10 · Faraday's law

Completed notes | Notes 2/3

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)

B10 · Faraday's law

Completed notes | 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 = **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.

B10 · Worked practice

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.

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

B11 · Faraday's law with the unknown chosen

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.

B11 · Faraday's law with the unknown chosen

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.

B12 · Industrial applications and corrosion protection

Completed notes | Notes 1/3

Explain rusting as a redox cell, choose a method to protect iron using the reduction table, describe plating, refining, electrowinning and aluminium production, and judge claims with evidence.

1. Rusting and barriers

Rusting is a tiny cell

Rusting is just redox, like B1. Corrosion is the slow wearing away of a metal by a **redox** reaction. At one spot, iron loses electrons: $Fe \rightarrow Fe^{2+} + 2e^-$.

The electrons travel through the iron to another spot. There, oxygen and water gain them: $O_2 + 2H_2O + 4e^- \rightarrow$ **$4OH^-$**. So iron needs both **water** and **oxygen** to rust.

The higher table value is the cathode, so oxygen (+0.40 V) is reduced. E°cell = cathode − anode = 0.40 − (−0.45) = **+0.85 V**.

A positive value says rusting is **spontaneous**, so it happens by itself. It does not say how **fast** rusting is. Salt water speeds it up because its ions carry charge.

On the diploma: to find what is oxidized, look for the particle that loses electrons. A wrong option flips the sign or adds the numbers.

Barriers

A barrier is a coating that keeps water and oxygen off the iron. Paint, oil and plastic are barriers. A barrier does not take part in the redox reaction.

It works only while whole. A deep **scratch** lets water and oxygen reach the iron again, and rusting can creep under the paint.

A metal coating is a barrier too. At a scratch, the metal lower on the table is oxidized first, because it gives up electrons more easily.

Tin (−0.14 V) is above iron (−0.45 V), so at a scratch the **iron** corrodes. Zinc (−0.76 V) is below iron, so the **zinc** corrodes and protects the iron.

On the diploma: predict by comparing the two table values and naming the lower one. The wrong option picks the higher value.

Rusting: $E°_{cell} = 0.40 - (-0.45) = +0.85$ V

Tin-plated steel, scratched (Sn −0.14 V). Corrodes: iron

Zinc-plated steel, scratched (Zn −0.76 V). Corrodes: zinc

Mg block on iron: $-0.45 - (-2.37) = +1.92$ V; anode is Mg

B12 · Industrial applications and corrosion protection

Completed notes | Notes 2/3

2. Sacrificial anodes and power supplies

Choose a sacrificial metal

A sacrificial anode is a block of metal that gives up electrons more easily than iron. Join it to the iron, and the block is oxidized instead.

So the block must be **below** iron on the table. That makes it the stronger reducing agent, which means it loses electrons more easily. It is the **anode** and the iron is the cathode.

Compare the metals with protected minus sacrificial. For Mg: $-0.45 - (-2.37) =$ **+1.92 V**. For Zn: $-0.45 - (-0.76) =$ **+0.31 V**. Positive means the metal works.

These are comparison numbers, not the real cell voltage. In the real cell, oxygen (+0.40 V) is reduced at the iron. With magnesium that is $0.40 - (-2.37) = +2.77$ V.

Copper (+0.34 V) is above iron, so it **fails**. The iron would become the anode and corrode faster. On the diploma: the wrong options are the metals above iron.

The block needs an electrical **connection** to the iron. It also needs an electrolyte, a liquid with dissolved ions such as wet soil or sea water. It is used up, so it must be **replaced**.

A larger number like +1.92 V shows magnesium is the stronger reducing agent. It does not prove lower cost or a longer **life**.

Use a power supply

A power supply pushes electrons around a circuit. A terminal is a connection point on it. Join the pipeline to the **negative** terminal.

Electrons are pushed onto the pipeline, so it is the **cathode** of an electrolytic cell. That is a cell that uses electrical energy to force a reaction.

A separate inert anode goes on the positive terminal. Inert means it does not react or dissolve. Oxidation happens there, not at the iron.

Both methods keep the protected metal as the cathode, so it is not oxidized. A sacrificial anode needs new blocks. A power supply needs **electricity**.

On the diploma: the protected metal always goes on the negative terminal. The wrong option is the positive one.

B12 · Industrial applications and corrosion protection

Completed notes | Notes 3/3

3. Making, cleaning and coating metals

Electrolytic processes

Electroplating coats an object with a thin layer of metal. The object is the **cathode**, because metal ions are reduced onto it: $Ag^+ + e^- \rightarrow$ **Ag**.

Refining purifies a metal. The impure metal is the **anode** and dissolves: $Cu \rightarrow Cu^{2+} + 2e^-$. Pure metal forms at the cathode.

Electrowinning recovers a metal from a solution of its ions. The anode is **inert** and does not dissolve, so the metal comes from the solution.

The Faraday chain works for all three processes. On the diploma: a common wrong choice says the anode dissolves in electrowinning.

Watch: Plating a copper ring with silver (Tyler DeWitt, “Electroplating”, 1:30)

Watch: The whole plating cell in one picture (Tyler DeWitt, “Electroplating Part 2”, 7:22)

Aluminium

Al^{3+}/Al is −1.66 V, very low on the table. In a water solution, the higher value is reduced first. Water is −0.83 V, so **hydrogen** forms at the cathode, not aluminium.

The fix is to remove the water. Aluminium oxide is dissolved in **molten** cryolite, a melted mineral with no water. Then Al^{3+} (−1.66 V) beats Na^+ (−2.71 V): $Al^{3+} + 3e^- \rightarrow$ **Al**.

The oxide ions are oxidized at a carbon anode. The oxygen reacts with the carbon, so the anode is **used up**.

A table value shows that a method **works**. It does not prove lower cost or a longer life.

Redox in living things

Oxidation numbers help us follow electrons. In cellular respiration, $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$, the average oxidation number of carbon rises from 0 to **+4**. So glucose is **oxidized**.

Six carbons each lose 4 electrons, so 24 are lost. Oxygen falls from 0 to −2, so O_2 is **reduced** and gains the same 24 electrons.

Oxygen is the oxidizing agent and glucose is the **reducing** agent. Respiration is spontaneous, so it releases energy, like a voltaic cell.

Photosynthesis is the **reverse**. It is not spontaneous by itself, so **light** supplies the energy, like a power supply in an electrolytic cell.

On the diploma: the particle whose oxidation number goes down is the one reduced.

B12 · Worked practice

Industrial applications and corrosion protection | 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.

B12-P01 · Easy · Iron rusts in damp air. Write the oxidation and reduction half-reactions and find E°cell (O_2/OH^- +0.40 V; Fe^{2+}/Fe −0.45 V).

1. Oxidation: $Fe \rightarrow Fe^{2+} + 2e^-$. Reduction: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$.

2. $E°cell = 0.40 - (-0.45) = +0.85$ V, so rusting is spontaneous.

B12-P02 · Easy · A student says that a positive E°cell of +0.85 V means iron rusts quickly. Explain what the value shows and what it does not show.

1. A positive value says that the reaction is spontaneous under standard conditions.

2. It does not say how fast the reaction goes. Salt water and other conditions change the rate.

B12 · Worked practice

Industrial applications and corrosion protection | Questions 3-4 of 12

B12-P03 · Easy · Explain how a layer of paint protects iron, and give one way that it can fail.

1. Paint is a barrier. It keeps water and oxygen away from the iron.

2. A deep scratch lets water and oxygen reach the iron at that spot, so rusting can start there.

B12-P04 · Easy · Tin-plated steel is scratched (Sn^{2+}/Sn −0.14 V; Fe^{2+}/Fe −0.45 V). Which metal corrodes at the scratch, and why?

1. Iron is below tin on the table, so iron is the stronger reducing agent.

2. The iron is oxidized at the scratch, so it corrodes, and the tin is the cathode.

B12 · Worked practice

Industrial applications and corrosion protection | Questions 5-6 of 12

B12-P05 · Medium · Zinc-plated steel is scratched (Zn^{2+}/Zn −0.76 V; Fe^{2+}/Fe −0.45 V). Which metal corrodes, and why does the iron stay protected?

1. Zinc is below iron on the table, so zinc is the stronger reducing agent and is oxidized first.

2. The iron is the cathode. The zinc near the scratch is used up instead of the iron.

B12-P06 · Medium · A steel pipe in wet soil is joined to a magnesium block (Mg^{2+}/Mg −2.37 V; Fe^{2+}/Fe −0.45 V). Name the anode and find protected minus sacrificial.

1. Magnesium is the anode, because it is oxidized: $Mg \rightarrow Mg^{2+} + 2e^{-}$. The steel is the cathode.

2. Protected minus sacrificial = $-0.45 - (-2.37) = +1.92$ V. This is a comparison number, not the real cell voltage.

B12 · Worked practice

Industrial applications and corrosion protection | Questions 7-8 of 12

B12-P07 · Medium · A plan uses copper blocks to protect a steel pipe (Cu^{2+}/Cu +0.34 V). Will the plan work? Explain with the table.

1. Copper is above iron on the table, so iron is the stronger reducing agent.

2. The iron would be the anode and corrode faster. Zinc or magnesium would work.

B12-P08 · Medium · The metals to choose from for protecting a steel hull are Cu (+0.34 V), Zn (−0.76 V), Sn (−0.14 V) and Pb (−0.13 V), with Fe^{2+}/Fe at −0.45 V. Choose the metal and give protected minus sacrificial.

1. Only zinc is below iron on the table, so only zinc can be the sacrificial metal.

2. Protected minus sacrificial = −0.45 − (−0.76) = +0.31 V, so zinc is oxidized first.

B12 · Worked practice

Industrial applications and corrosion protection | Questions 9-10 of 12

B12-P09 · Challenging · A sacrificial block must be replaced from time to time. Explain why, and name one other thing that the block needs to work.

1. The block is oxidized, so its atoms leave as ions and it is used up.

2. It needs an electrical connection to the iron and an electrolyte, such as wet soil or sea water.

B12-P10 · Challenging · A buried steel pipeline is protected with a power supply. State the terminal that the pipeline is joined to and the electrode where oxidation happens.

1. The pipeline is joined to the negative terminal, so it is the cathode.

2. Oxidation happens at a separate inert anode that is joined to the positive terminal.

B12 · Worked practice

Industrial applications and corrosion protection | Questions 11-12 of 12

B12-P11 · Challenging · A steel spoon is silver plated. State whether the spoon is the anode or the cathode, and write the half-reaction at the spoon.

1. The spoon is the cathode, because the silver ions are reduced onto it.

2. The half-reaction is $Ag^{+} + e^{-} \rightarrow Ag$.

B12-P12 · Challenging · A student tries to make aluminium by electrolysis of aqueous Al^{3+} with inert electrodes (Al^{3+}/Al −1.66 V; water −0.83 V). State what forms at the cathode and how aluminium is made in industry.

1. Water is reduced because −0.83 V is higher than −1.66 V, so hydrogen forms at the cathode.

2. Industry uses molten aluminium oxide in a salt that has no water, so $Al^{3+} + 3e^{-} \rightarrow Al$. The carbon anode is used up.

B13 · Batteries, fuel cells and chlorine for water

Completed notes | Notes 1/3

Explain batteries as voltaic cells and write the reactions of lead-acid, alkaline and fuel cells. Explain chlorine production for water and judge a cell using evidence.

1. Batteries

A battery is one or more voltaic cells

A voltaic cell has an anode, a cathode and an electrolyte, which is a liquid with dissolved ions. A battery is one or more of them.

Cells joined in a row **add** their voltages, so more cells means more voltage.

A **primary** cell is used once, because its reaction cannot easily be reversed. A **secondary** cell can be recharged.

During discharge a battery is a **voltaic** cell. The reaction runs by itself and gives electricity.

During recharge a power supply forces the reverse reaction, so it is an **electrolytic** cell (B8). The electrode that was oxidized is now reduced, so the names anode and cathode swap.

On the diploma: to name the anode, ask where oxidation is happening in that direction. A battery can be one cell or several cells in a row.

The lead-acid cell

The booklet has two lead rows: $PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$ (+1.69 V) and $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$ (−0.36 V).

PbO_2 has the higher value, +1.69 V, so it is the **cathode**. The lower row runs backward at the anode: $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$. So lead is the **anode**.

E°cell = 1.69 − (−0.36) = **+2.05 V**. One cell is not enough for a car, and voltages add. Six cells in a row give 6 × 2.05 = **12.3 V**.

Both halves move two electrons, so no multiplier is needed. The cell reaction is $PbO_2 + Pb + 2SO_4^{2-} + 4H^+ \rightarrow$ **$2PbSO_4$** $+ 2H_2O$. Atom check: Pb 2, S 2, O 10, H 4, charge 0.

H^+ and SO_4^{2-} are reactants, so they get used up and the acid gets **weaker**. $PbSO_4$ is a solid that forms on both electrodes, so both electrodes **gain** mass.

We cannot use part of a cell. If each cell gives 2.05 V and a battery needs 7.0 V, then 7.0 ÷ 2.05 = 3.41, so round up to **4** cells.

On the diploma: reverse the lower half-reaction but keep its E° as printed, and never multiply E° by a coefficient. Common wrong choices forget the six cells or say the

anode loses mass.

PbO_2 (+1.69 V) and $PbSO_4$/Pb (−0.36 V). Cathode: PbO_2; anode: Pb

$E°cell = 1.69 - (-0.36) = +2.05$ V

Six cells in a row: $6 \times 2.05 = 12.3$ V

Discharge: both electrodes gain mass as $PbSO_4$ forms.

B13 · Batteries, fuel cells and chlorine for water

Completed notes | Notes 2/3

2. Alkaline cells and fuel cells

The alkaline cell

An alkaline cell, like a flashlight cell, is a primary cell. Its half-reactions are given, so we do not need the table.

The anode is **zinc**, which is oxidized because it loses electrons: $Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$. The cathode is **$MnO_2$**, which is reduced: $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$.

Add the halves and cancel the electrons, water and hydroxide ions. The overall reaction is $Zn + 2MnO_2 \rightarrow ZnO +$ **Mn_2O_3**. It is a **primary** cell with a given voltage of about 1.5 V.

Used cells contain **metals**, like zinc and manganese compounds. In the garbage they can harm soil and water, so they should be recycled or disposed of properly.

On the diploma: when half-reactions are given, use them and do not search the booklet. An Evaluate item about batteries expects a benefit and a disposal concern.

The hydrogen fuel cell

Hydrogen is **oxidized** at the anode: $2H_2 \rightarrow 4H^+ + 4e^-$. Oxygen is **reduced** at the cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$.

The electrons travel through the outside circuit, and that is the current we use. Hydrogen ions cross a **membrane** inside the cell to complete the circuit.

Both halves move four electrons, so they cancel. Overall: $2H_2 + O_2 \rightarrow$ **$2H_2O$**. $E°_{cell} = 1.23 - 0.00 =$ **+1.23 V**, and we do not multiply E° by the 2.

A fuel cell is fed **fuel** all the time, so it keeps running. A battery stores a fixed amount of **reactants** and runs down.

On the diploma: the particle that is oxidized is the one that loses electrons. Do not multiply E° by the coefficient you used to balance electrons.

B13 · Batteries, fuel cells and chlorine for water

Completed notes | Notes 3/3

3. Chlorine for water and judging cells

Chlorine for a pool

A pool system passes current through NaCl(aq). The cathode makes **hydrogen**: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$. The anode makes **chlorine**: $2Cl^- \rightarrow Cl_2 + 2e^-$.

The table says water should be oxidized, but chlorine forms because of the chloride **anomaly** from B9.

The Cl_2 then reacts with hydroxide: $Cl_2 + 2OH^- \rightarrow Cl^- + OCl^- + H_2O$. One chlorine goes from 0 to **−1**, so it is reduced. The other goes from 0 to **+1**, so it is oxidized.

The same element is both oxidized and reduced, which we call a **disproportionation**.

The ion **OCl^-** cleans the water, because it is an oxidizing agent that oxidizes bacteria.

On the diploma: write the oxidation number of chlorine on each side before you name the change. Hydrogen is at the cathode and chlorine at the anode.

Judge a cell

Start with the **job**. Then compare voltage, mass, lifetime, recharging, cost and **disposal** as separate pieces of evidence.

A higher voltage does not prove a **longer** life or a lower **cost**. A primary cell is cheap and keeps its charge in storage, so it suits a flashlight that is used rarely.

A fuel cell runs as long as fuel arrives. But a claim of no emissions only covers the cell itself, so we need evidence about how the **fuel** is made and delivered.

On the diploma: a common wrong choice says the highest-voltage cell is best. Justify your choice with one advantage and one limit.

Extracting metals long ago

Early peoples heated copper ore with charcoal long before anyone knew about electrons: $2CuO + C \rightarrow 2Cu + CO_2$. Copper goes from +2 to 0, so it is **reduced**. Carbon goes from 0 to +4, so it is **oxidized**.

Carbon gave electrons to copper(II), so carbon is the **reducing** agent. It is the one whose oxidation number goes up.

Trial and error came first, and the redox **model** explained it later. The model also showed that metals low on the table, like aluminium, have ions that accept electrons so poorly that carbon cannot reduce them. Those need electrolysis (B12).

On the diploma: for a science and technology question, give one example in each direction.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | 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.

B13-P01 · Easy · Explain the difference between a primary cell and a secondary cell, and give one example of each from this lesson.

1. A primary cell is used once and then disposed of or recycled. An alkaline flashlight cell is an example.

2. A secondary cell can be recharged. A lead-acid car battery is an example.

B13-P02 · Easy · A lead-acid car battery is recharged. State what kind of cell it is during recharging, and what the power supply does.

1. During recharging the battery is an electrolytic cell.

2. The power supply forces the reaction backward, so $PbSO_4$ changes back to Pb and PbO_2.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 3-4 of 12

B13-P03 · Easy · Given $PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$ (+1.69 V) and $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$ (−0.36 V). Name the cathode and the anode for a discharging lead-acid cell.

1. The higher value, +1.69 V, is reduced at the cathode, so PbO_2 is the cathode.

2. The lead is oxidized at the anode: $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$.

B13-P04 · Easy · Write the cell reaction for the discharging lead-acid cell and find E°cell.

1. Both halves move two electrons, so add them without a multiplier.

2. $PbO_2 + Pb + 2SO_4^{2-} + 4H^+ \rightarrow 2PbSO_4 + 2H_2O$. E°cell = 1.69 − (−0.36) = +2.05 V.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 5-6 of 12

B13-P05 · Medium · A car battery has six lead-acid cells in a row. Find its voltage and explain why cells in a row are used.

1. The voltages add: 6 × 2.05 = 12.3 V.

2. One cell gives only 2.05 V, so many cells are joined to get a voltage large enough for the car.

B13-P06 · Medium · Describe what happens to the acid and to the mass of each electrode as a lead-acid battery discharges.

1. Hydrogen ions and sulfate ions are used up, so the acid gets weaker.

2. Solid $PbSO_4$ forms on both electrodes, so each electrode gains mass.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 7-8 of 12

B13-P07 · Medium · Given the alkaline cell half-reactions $Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$ and $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$. Add them and write the overall reaction.

1. Each half moves two electrons. Cancel the electrons, water and hydroxide ions.

2. The overall reaction is $Zn + 2MnO_2 \rightarrow ZnO + Mn_2O_3$.

B13-P08 · Medium · Why should used alkaline cells be recycled or disposed of properly instead of being thrown into the garbage?

1. Used cells contain metals and other chemicals.

2. These can harm the environment if they leak into soil or water, so recycling recovers the metals and reduces the harm.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 9-10 of 12

B13-P09 · Challenging · Given $H_2 \rightarrow 2H^+ + 2e^-$ (0.00 V) and $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (+1.23 V). Write the overall fuel cell reaction and find E°cell.

1. Multiply the hydrogen half by 2 so that four electrons match.

2. $2H_2 + O_2 \rightarrow 2H_2O$. E°cell = 1.23 − 0.00 = +1.23 V.

B13-P10 · Challenging · A company says that its fuel cell makes only water, so its total emissions are zero. Evaluate the claim.

1. The claim is about the cell where it runs. It does not include the making and delivery of the hydrogen.

2. We need evidence about how the hydrogen is made, squeezed into tanks and moved. We cannot accept or reject the claim without it.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 11-12 of 12

B13-P11 · Challenging · A flashlight is used only now and then. Compare an alkaline cell and a rechargeable cell for this job, and name the evidence that you would use.

1. An alkaline primary cell is cheap and keeps its charge when stored, so it suits a device that is used rarely.

2. A rechargeable cell costs more at the start and pays off only when it is recharged many times. We compare cost, lifetime and use, not voltage alone.

B13-P12 · Challenging · Salt water is electrolyzed in a pool system. Name the product at each electrode, and show with oxidation numbers that $Cl_2 + 2OH^- \rightarrow Cl^- + OCl^- + H_2O$ is a disproportionation.

1. Hydrogen forms at the cathode and chlorine forms at the anode.

2. In Cl_2 the oxidation number is 0. In Cl^- it is -1 and in OCl^- it is $+1$, so chlorine is both reduced and oxidized.

B14 · Investigation: plan and predict

Completed notes | Notes 1/2

Write a testable question, draw a labelled cell design, predict the anode and voltage from the table, plan safe handling and plan a battery.

1. Question, variables and design

A testable question

We have used the table to predict voltages. Now we test the predictions. A testable question links one thing we change to one thing we measure.

The variable you change is the **independent** variable. The variable you measure is the **dependent** variable. Here we change the metal and measure the voltage.

A fair test keeps everything else the same. These are the controlled variables: the **concentration**, temperature, electrode size and salt bridge.

The table values assume **1.0 mol/L** solutions at standard conditions. If a control changed, we could not tell what caused a change in voltage.

On the diploma: the independent variable is what you change, and the dependent variable is what you measure.

A labelled design

A half-cell is a metal strip, called an electrode, in a solution of its **ions**. The second half-cell is always copper in copper(II) sulfate.

A **salt bridge** lets ions move between the half-cells. This keeps the charge balanced and completes the circuit. A voltmeter joins the strips through wires.

Label the **anode**, the cathode, the direction of electron flow and the concentrations. Electrons flow in the wire from the anode to the cathode.

On the diploma: a labelled diagram needs the salt bridge, the solutions with concentrations, the wires and the meter.

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

B14 · Investigation: plan and predict

Completed notes | Notes 2/2

2. Predict, stay safe and plan a battery

Predict before measuring

The metal with the **lower** reduction value is the anode, because that metal is oxidized. The higher value is the cathode.

Zn/Cu: zinc (−0.76 V) is lower, so zinc is the anode. E°cell = cathode − anode = 0.34 − (−0.76) = **+1.10 V**. Dropping the minus sign gives a classic wrong answer.

Ag/Cu: silver (+0.80 V) has the higher value, so copper is the **anode**. E°cell = 0.80 − 0.34 = **+0.46 V**. Copper is not always the cathode.

When you double a half-reaction to balance electrons, you do not **double** its table value. The voltage does not depend on the number of copies.

A prediction is not an **observation**. Only a measurement shows what the cell does. Ask where a number came from: the table or a meter.

Safety and disposal

Read the **label** and the safety data sheet first. They are part of **WHMIS**, the system for labelling hazardous materials in workplaces.

Wear goggles, gloves and an apron. Silver nitrate stains skin. Copper(II) compounds harm life in water.

Collect solutions that contain metal **ions**. Do not pour them down the drain, because they can harm water life.

On the diploma: name the hazard from the label, then a matching precaution, and give a reason for each step.

Plan a battery

Cells in a row **add** their voltages. A device needs at least 3.0 V, and one Zn/Cu cell gives 1.10 V.

So 3.0 ÷ 1.10 = 2.7. We cannot use part of a cell, so round up and use **3** cells. They give 3 × 1.10 = 3.30 V. Two cells give only 2.20 V.

The voltage **falls** as the cell works, because the reactants get used up. So the plan must include a test over time.

A bigger table voltage does not always make a better battery. Magnesium gives 2.71 V with copper but reacts with water. On the diploma: name a practical fact such as safety, cost or lifetime.

B14 · Worked practice

Investigation: plan and predict | 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.

B14-P01 · Easy · Given the question "How does the metal in one half-cell change the cell voltage?" Name the independent variable and the dependent variable.

1. The independent variable is what you change, which is the metal.

2. The dependent variable is what you measure, which is the voltage.

B14-P02 · Easy · Name two variables that must be controlled when you compare metals joined to copper, and explain why.

1. Control the concentration (1.0 mol/L) and the temperature. Electrode size and the salt bridge also qualify.

2. If one of these changed, you could not tell whether the metal or the changed condition caused the change in voltage.

B14 · Worked practice

Investigation: plan and predict | Questions 3-4 of 12

B14-P03 · Easy · List the parts that a labelled diagram of a zinc and copper cell must show.

1. Two half-cells with the strips and solutions, the concentrations, the salt bridge, the wires and the voltmeter.

2. Label the anode, the cathode and the direction of electron flow in the wire.

B14-P04 · Easy · Explain what the salt bridge does in your design.

1. The salt bridge lets ions move between the two half-cells.

2. This keeps the charge balanced in each solution and completes the circuit, so the cell can keep running.

B14 · Worked practice

Investigation: plan and predict | Questions 5-6 of 12

B14-P05 · Medium · Given Fe^{2+}/Fe −0.45 V and Cu^{2+}/Cu +0.34 V. Predict the anode, the cell reaction and the voltage.

1. Iron has the lower value, so iron is the anode.

2. $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$. $E°cell = 0.34 - (-0.45) = +0.79$ V.

B14-P06 · Medium · Given Ag^{+}/Ag +0.80 V and Cu^{2+}/Cu +0.34 V. Which metal is the anode and what is the voltage?

1. Silver has the higher value, so silver is the cathode and copper is the anode.

2. $E°cell = 0.80 - 0.34 = +0.46$ V.

B14 · Worked practice

Investigation: plan and predict | Questions 7-8 of 12

B14-P07 · Medium · Given Ni^{2+}/Ni −0.26 V and Ag^{+}/Ag +0.80 V. Write the cell reaction and find the voltage. Explain why the silver table value is not doubled.

1. Nickel is the anode. Two silver ions take the two electrons: $Ni + 2Ag^{+} \rightarrow Ni^{2+} + 2Ag$.

2. $E°cell = 0.80 - (-0.26) = +1.06$ V. A table value does not depend on how many copies of the half-reaction are used.

B14-P08 · Medium · A student predicts +0.79 V and says this proves the measured voltage will be 0.79 V. Explain what is wrong.

1. A prediction is not an observation.

2. The table assumes standard conditions, so a real measurement can differ a little. Only a measurement shows what happens.

B14 · Worked practice

Investigation: plan and predict | Questions 9-10 of 12

B14-P09 · Challenging · Name two hazards or precautions for a plan that uses silver nitrate solution, and one disposal step.

1. Silver nitrate stains skin, so wear gloves, goggles and an apron. Read the label and the safety data sheet.

2. Collect the waste in a container. Do not pour it down the drain, because metal ions can harm water life.

B14-P10 · Challenging · A device needs at least 3.0 V. Using 1.10 V cells in a row, find how many cells are needed and the voltage of the battery.

1. The voltages add. 3.0 ÷ 1.10 = 2.7, and part of a cell cannot be used, so use 3 cells.

2. The battery voltage is 3 × 1.10 = 3.30 V. Two cells give only 2.20 V.

B14 · Worked practice

Investigation: plan and predict | Questions 11-12 of 12

B14-P11 · Challenging · Explain why a metal with a large cell voltage can still be a poor choice for a battery. Use magnesium as an example.

1. Magnesium with copper predicts +2.71 V.

2. Magnesium reacts with water, so it would be used up, and cost, mass and lifetime matter as well as voltage.

B14-P12 · Challenging · Given a plan and a prediction only. State what they can show and what they cannot show, and why a real measurement may differ.

1. A plan and a prediction show what you expect. They are not data.

2. Real strips, salt bridges and meters have small errors, and the table applies only at standard conditions.

B15 · Investigation: analyze and explain

Completed notes | Notes 1/2

Put together given evidence, compare predictions with observations, explain differences, identify products and judge a battery design.

1. Evidence and comparison

Put together the evidence

In B14 we predicted each voltage. Now we look at results. This lesson uses a **supplied** sample set that we did not measure.

Pairs are predicted/measured: Zn 1.10/1.07 V, Fe 0.79/0.76 V, Ni 0.60/0.57 V, Ag 0.46/0.43 V. Each row is one cell.

A good table has a title, **units** in the headings, the predicted column and the measured column. Note the controlled conditions too.

The measured column holds **observations**, because those numbers came from a voltmeter. The predicted column holds values from the reduction table.

On the diploma: a common wrong choice mixes up observation and prediction. Ask which column came from a measurement.

Compare

Difference = predicted − **measured**. Each cell here is **0.03** V lower than the prediction. That is a pattern with a direction and a size.

The order also matches, with zinc highest and silver lowest. So the data **support** the table. We say support, not prove.

A percent difference compares the gap with the prediction. For zinc, 0.03 ÷ 1.10 × 100 = **2.7** %. A small gap means close agreement.

On the diploma: describe a pattern with its direction and size, with the unit. Do not just say the numbers are close.

B15 · Investigation: analyze and explain

Completed notes | Notes 2/2

2. Explain, identify and evaluate

Explain a difference

Three possible causes fit a constant gap: the salt bridge adds **resistance**, the strips have **oxide** coatings, or the concentration is not exactly 1.0 mol/L.

Temperature is another, because the table assumes 25 °C. A good report records the room temperature and the concentration.

To test a cause, change **one** thing and measure again. For example, sand the strips and see if the voltage rises toward the prediction.

Until the test is done, the cause is **possible**, not proven. On the diploma: never claim a cause that was not tested.

Identify products

In the Zn/Cu cell, the zinc strip loses mass, so zinc is oxidized. The copper strip gains a **reddish-brown** coating, so copper ions are reduced.

An observation is what we see or measure. An inference is a conclusion drawn from it. Zn^{2+} is colourless, so it is an **inference**.

In the Ag/Cu cell, copper is the anode. The copper strip loses mass, and the copper solution is already blue, so it gets **deeper** blue as Cu^{2+} forms.

A conclusion has a **claim**, the evidence and an explanation. Name a limitation and an **improvement**, such as three trials averaged.

On the diploma: to identify a product, name the observation that supports it.

Battery design and disposal

Three cells of 1.07 V give $3 \times 1.07 =$ **3.21 V**. Judge voltage, mass, cost, lifetime and **leaks**, because voltage is only one part of the evidence.

Used batteries contain metals that can harm soil and water, so they are collected for **recycling**. Recycling has benefits and costs, like the energy it needs.

A hydrogen fuel cell gives about 1.23 V per cell. A fair presentation gives purpose, reactions, benefits and **limits**, and states the source of its evidence.

On the diploma: to evaluate a technology, name evidence that supports it and a limit that remains.

B15 · Worked practice

Investigation: analyze and explain | 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.

B15-P01 · Easy · Name three things that a good data table includes.

1. A title, headings with units, and the predicted values beside the measured values.

2. The controlled conditions should also be noted so that a reader can judge the test.

B15-P02 · Easy · Which column of the table holds observations and which holds predictions? Explain.

1. The measured column holds observations, because they come from measuring.

2. The predicted column holds values from the reduction table.

B15 · Worked practice

Investigation: analyze and explain | Questions 3-4 of 12

B15-P03 · Easy · Given a given result for zinc and copper: predicted 1.10 V, measured 1.07 V. Find the difference.

1. The difference is predicted minus measured.

2. 1.10 − 1.07 = 0.03 V, so the measured value is 0.03 V lower.

B15-P04 · Easy · All four given cells are 0.03 V below the prediction and the order matches. Does the data support the table? Give a reason.

1. Yes. The order of the voltages is the same as predicted.

2. Each value is within 0.03 V of the prediction, so the table predicts the order and the size well.

B15 · Worked practice

Investigation: analyze and explain | Questions 5-6 of 12

B15-P05 · Medium · Give two possible explanations for a small constant gap and say how to test one of them.

1. The salt bridge adds resistance, and the strips may have an oxide coating.

2. Sand the strips, measure again and see if the voltage moves closer to the prediction. Change only one thing.

B15-P06 · Medium · Explain why a possible cause should not be stated as proven before a test.

1. A possible cause fits the evidence but has not been tested.

2. Only a test that changes one thing shows if that thing causes the gap.

B15 · Worked practice

Investigation: analyze and explain | Questions 7-8 of 12

B15-P07 · Medium · In a zinc and copper cell the zinc strip loses mass and the copper strip gains a reddish-brown coating. State two observations and the half-reactions they support.

1. Observations: the zinc loses mass and the copper gains a coating.

2. $Zn \rightarrow Zn^{2+} + 2e^{-}$ and $Cu^{2+} + 2e^{-} \rightarrow Cu$.

B15-P08 · Medium · $Zn^{2+}(aq)$ is colourless. Is its formation an observation or an inference? Explain.

1. It is an inference, not an observation.

2. You cannot see the ions. You reason that they form because the zinc strip loses mass.

B15 · Worked practice

Investigation: analyze and explain | Questions 9-10 of 12

B15-P09 · Challenging · Given a given nickel and copper result: predicted 0.60 V, measured 0.57 V. Write a short conclusion with a claim, evidence and one possible explanation.

1. Claim: the result agrees with the table. Evidence: 0.57 V against 0.60 V and the strip changes.

2. Explanation: the salt bridge or an oxide coating could lower the voltage a little.

B15-P10 · Challenging · A design uses one trial for each cell. Give one limitation and one improvement.

1. One trial cannot show how much the measurement varies.

2. Repeat three trials and take the average, and clean the strips before each trial.

B15 · Worked practice

Investigation: analyze and explain | Questions 11-12 of 12

B15-P11 · Challenging · Four cells of 1.07 V are joined in a row. Find the battery voltage and name one disposal step for the used battery.

1. The voltages add: 4 × 1.07 = 4.28 V.

2. Collect the used battery for recycling. Do not put it in the garbage, because it contains metals that can harm soil and water.

B15-P12 · Challenging · Name two parts that a short fuel cell presentation must include, and say why the source of evidence should be stated.

1. Include any two of the purpose, the half-reactions, the benefits and the limits.

2. The source lets a reader check the claims. A claim without evidence cannot be judged.

B16 · Unit B review: choose the method

Completed notes | Notes 1/2

Sort a problem, choose between the table, Faraday and titration, and apply Unit B ideas to corrosion, plating, refining and batteries.

1. Choose the method

Sort the problem

Choose the method before you touch a calculator. Voltage, spontaneity or the anode: use the **reduction** table. Mass, current or time at an electrode: use **Faraday**'s law.

A concentration from a volume of a reactant: use a redox **titration**. It finds a concentration by reacting a sample with a solution of known concentration.

A positive E°cell means the reaction runs by itself, so it can be a **voltaic** cell. A negative E°cell needs a power supply, so the cell is **electrolytic**.

Either way, oxidation happens at the **anode**. The electrode names follow the reaction, not the cell type.

On the diploma: many wrong options use the right numbers in the wrong method, so name the method first.

Watch: Is this equation redox or not? (Alberta Chemistry Teacher, "Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4", 18:39)

Watch: Spotting disproportionation (Alberta Chemistry Teacher, "Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4", 20:04)

Methods in order

Table: the anode is the **lower** value, E°cell = cathode − anode. Do not **multiply** E° values by coefficients, because voltage does not depend on the number of copies.

Faraday: q = **It** with the time in seconds, then $n(e^-) = q \div 9.65 \times 10^4$, then the mole **ratio** from the half-reaction, then the mass.

Titration: balance the **equation**, find the amount of the known reactant, use the ratio, then divide by the volume in **litres**. Divide millilitres by 1000.

The endpoint is the moment the reaction is just complete. With permanganate, it is the first lasting pale pink, so no indicator is needed.

On the diploma: a wrong mole ratio is a common wrong option, so read it from the half-reaction or the balanced equation.

Watch: Trends in the reduction table (Alberta Chemistry Teacher, "Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4", 23:36)

Watch: Patterns: metals, halogens and multivalent ions (Alberta Chemistry Teacher, "Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4", 24:21)

B16 · Unit B review: choose the method

Completed notes | Notes 2/2

2. Apply the ideas

Corrosion and industry

Magnesium (−2.37 V) is below iron (−0.45 V), so it is a stronger reducing agent. It is the **sacrificial** anode that is oxidized, and the iron is protected as the cathode.

A power supply protects iron too, by making it the cathode. Both methods stop the iron from being oxidized.

In electroplating the object is the **cathode**. In refining the impure metal is the **anode**. Oxidation is always at the anode, reduction at the cathode.

Judge a battery by voltage, mass, cost, lifetime and **disposal**. On the diploma: a wrong electrode is a common wrong choice, so label the anode and cathode first.

B16 · Worked practice

Unit B review: choose the method | 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.

B16-P01 · Easy · Which method do you use for a question that gives 2.00 A for 30.0 min and asks for a mass of copper? Explain.

1. Use Faraday’s law for a mass at an electrode.

2. What is given are a current and a time, and the question asks for a mass at an electrode.

B16-P02 · Easy · Given a calculated E°cell of −0.50 V. State the type of cell that can make the reaction happen and what it needs.

1. A negative E°cell means the reaction is not spontaneous.

2. It needs an electrolytic cell and a power supply that provides more than 0.50 V.

B16 · Worked practice

Unit B review: choose the method | Questions 3-4 of 12

B16-P03 · Easy · Given Al^{3+}/Al −1.66 V and Ag^{+}/Ag +0.80 V. Write the cell reaction.

1. Aluminium is the anode and gives three electrons. Each silver ion takes one, so triple the silver half.

2. The cell reaction is $Al + 3Ag^{+} \rightarrow Al^{3+} + 3Ag$.

B16-P04 · Easy · Find the voltage of the aluminium and silver cell, and explain why the silver value is not multiplied by three.

1. $E°cell = 0.80 - (-1.66) = +2.46$ V.

2. A table value does not depend on the number of copies of the half-reaction.

B16 · Worked practice

Unit B review: choose the method | Questions 5-6 of 12

B16-P05 · Medium · A current of 3.00 A runs for 25.0 min with a silver cathode in Ag^+ solution. Find the charge and the amount of electrons.

1. Time = 25.0 × 60 = 1500 s, so q = 3.00 × 1500 = 4500 C.

2. $n(e^-) = 4500 \div (9.65 \times 10^4) = 0.0466$ mol.

B16-P06 · Medium · Continue the last question and find the mass of silver deposited. M(Ag) = 107.87 g/mol.

1. The ratio is 1 to 1, so n(Ag) = 0.0466 mol.

2. m = 0.0466 × 107.87 = 5.03 g.

B16 · Worked practice

Unit B review: choose the method | Questions 7-8 of 12

B16-P07 · Medium · Given 4.00 A for 20.0 min through $Cu^{2+}(aq)$ with a copper cathode. Find the mass of copper deposited. M(Cu) = 63.55 g/mol.

1. q = 4.00 × 1200 = 4800 C. $n(e^-) = 4800 \div (9.65 \times 10^4) = 0.04974$ mol.

2. Two electrons deposit one copper atom, so n(Cu) = 0.02487 mol and m = 1.58 g.

B16-P08 · Medium · Given $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. If 20.00 mL of 0.0150 mol/L MnO_4^- reacts with 25.00 mL of $Fe^{2+}(aq)$, find $[Fe^{2+}]$.

1. $n(MnO_4^-) = 0.02000 \times 0.0150 = 3.00 \times 10^{-4}$ mol. The ratio is 5 to 1, so $n(Fe^{2+}) = 1.50 \times 10^{-3}$ mol.

2. $[Fe^{2+}] = 1.50 \times 10^{-3} \div 0.02500 = 0.0600$ mol/L.

B16 · Worked practice

Unit B review: choose the method | Questions 9-10 of 12

B16-P09 · Challenging · Describe the end point of the permanganate titration and say why no indicator is added.

1. The end point is the first lasting pale pink.

2. Permanganate is purple and fades as it reacts, so it is its own indicator.

B16-P10 · Challenging · Explain how a magnesium block protects a steel pipe in soil. Name the role of the pipe.

1. Magnesium has a lower reduction value, so it is oxidized instead of iron.

2. The magnesium is the anode and is used up, and the pipe is the cathode.

B16 · Worked practice

Unit B review: choose the method | Questions 11-12 of 12

B16-P11 · Challenging · A spoon is plated with silver using a silver nitrate solution. Name the electrode for the spoon, the half-reaction there and the power supply connection.

1. The spoon is the cathode, where silver ions are reduced: $Ag^{+} + e^{-} \rightarrow Ag$.

2. The spoon is connected to the negative terminal of the power supply.

B16-P12 · Challenging · Choose the method and solve: 2.00 A runs for 30.0 min through molten NaCl. Find the mass of sodium at the cathode. M(Na) = 22.99 g/mol.

1. Use Faraday. $q = 2.00 \times 1800 = 3600$ C and $n(e^{-}) = 3600 \div (9.65 \times 10^{4}) = 0.0373$ mol.

2. The ratio is 1 to 1, so $m = 0.0373 \times 22.99 = 0.858$ g.