B16 · Unit B review: choose the method

Completed notes

Goals: Choose the method a Unit B question needs. Use the table, Faraday’s law and a titration in turn. Explain corrosion protection, plating and bleaching with redox ideas.

Slide 1 of 8 · Sort the problem before you solve it

Choose the method before you touch a calculator. If the question asks about voltage, spontaneity or the anode, use the **E° table**. If it gives current and time, or asks for a mass at an electrode, use **Faraday’s law**.

If it asks for a concentration from the volume of a reactant, use a **redox titration**. What is given and the **unknown** point you to the method.

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

In both cell types, **oxidation** happens at the anode and **reduction** happens at the cathode. Electrode names follow the reaction, not the cell type.

Example. Each Unit B problem asks about a cell, a charge or a titration. Choose the method for each. (a) Find the voltage of a zinc and silver cell. (b) Find the mass of silver from 3.00 A for 25.0 min. (c) Find $[Fe^{2+}]$ from the volume of MnO_4^- solution needed.

1. (a) Voltage: use the E° table.

2. (b) Current and time, and a mass at an electrode: use Faraday’s law.

3. (c) A concentration from a volume of reactant: use a redox titration.

Try it. Choose the method for each. (a) Find the mass of copper from 2.00 A for 10.0 min. (b) Find $[Fe^{2+}]$ from a permanganate titration. Then, for E°reaction = −0.50 V, name the cell type and the cathode reaction.

1. (a) Current and time, and a mass: use Faraday’s law.

2. (b) A concentration from a volume of reactant: use a redox titration.

3. E°reaction is negative, so the reaction needs a power supply and runs in an electrolytic cell. Reduction still happens at the cathode.

Slide 2 of 8 · Combine half-reactions

Aluminium has the **lower** E° (−1.66 V), so it is the anode and is oxidized: $Al \rightarrow Al^{3+} + \mathbf{3e^-}$. Silver is the cathode: $Ag^+ + e^- \rightarrow Ag$.

Each silver ion takes only 1 electron, so **triple** the silver half. Add the halves: $Al + 3Ag^+ \rightarrow \mathbf{Al^{3+} + 3Ag}$.

Check the balance. Al 1, Ag 3 and charge **+3** on each side. A balanced equation must match in atoms and in **charge**.

E°reaction = E°cathode − E°anode = 0.80 − (−1.66) = **+2.46 V**. Do not **multiply** the silver value by three. A positive value means the reaction is spontaneous.

Example. Find the cell reaction and voltage for aluminium and silver half-cells at 1.0 mol/L. Given $E°(Al^{3+}/Al) = -1.66$ V and $E°(Ag^+/Ag) = +0.80$ V.

1. Aluminium has the lower E°, so aluminium is the anode: $Al \rightarrow Al^{3+} + 3e^-$

2. Silver is the cathode. Triple it to match $3e^-$: $3Ag^+ + 3e^- \rightarrow 3Ag$

3. Add: $Al + 3Ag^+ \rightarrow Al^{3+} + 3Ag$

4. Check: Al 1, Ag 3, charge +3 on each side.

5. E°reaction = 0.80 − (−1.66) = +2.46 V, which is positive, so it is spontaneous.

Try it. Find the cell reaction and voltage for aluminium and copper half-cells at 1.0 mol/L. Given $E°(Al^{3+}/Al) = -1.66$ V and $E°(Cu^{2+}/Cu) = +0.34$ V.

1. Aluminium has the lower E°, so aluminium is the anode: $Al \rightarrow Al^{3+} + 3e^-$

2. Copper is the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$

3. Make $6e^-$ in each half: $2Al \rightarrow 2Al^{3+} + 6e^-$ and $3Cu^{2+} + 6e^- \rightarrow 3Cu$

4. Add: $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$. Check: Al 2, Cu 3, charge +6 on each side.

5. E°reaction = 0.34 − (−1.66) = +2.00 V, which is positive, so it is spontaneous.

Slide 3 of 8 · From current and time to a mass

Current is charge per second. Change minutes to **seconds** first. The charge is q = **It**, where I is the current in amperes and t is the time in seconds.

Divide the charge by Faraday's constant, **9.65 × 10⁴** C/mol e^-, to get the moles of electrons. Here 4500 ÷ (9.65 × 10⁴) = **0.04663** mol.

The mole ratio comes from the half-reaction. $Ag^+ + e^- \rightarrow Ag$ needs 1 mol of electrons per mole of metal, $Cu^{2+} + 2e^- \rightarrow Cu$ needs **2**, and $Al^{3+} + 3e^- \rightarrow Al$ needs **3**.

Mass = moles × molar mass. For silver, 0.04663 mol × 107.87 g/mol = **5.03 g**. A wrong mole ratio makes the mass wrong by that same factor.

Example. A current of 3.00 A runs for 25.0 min in a cell with a silver cathode in Ag^+ solution. Find the mass of silver deposited. Use F = 9.65 × 10⁴ C/mol e^- and M(Ag) = 107.87 g/mol.

1. t = 25.0 min × 60 = 1500 s

2. q = It = 3.00 A × 1500 s = 4500 C

3. $n(e^-)$ = 4500 C ÷ (9.65 × 10⁴) = 0.04663 mol

4. $Ag^+ + e^- \rightarrow Ag$ is 1 to 1, so n(Ag) = 0.04663 mol.

5. m = 0.04663 mol × 107.87 g/mol = 5.03 g

Try it. A current of 2.50 A runs for 18.0 min in a cell with a copper cathode in Cu^{2+} solution. Find the mass of copper deposited. Use F = 9.65 × 10⁴ C/mol e^- and M(Cu) = 63.55 g/mol.

1. t = 18.0 min × 60 = 1080 s

2. q = It = 2.50 A × 1080 s = 2700 C

3. $n(e^-)$ = 2700 C ÷ (9.65 × 10⁴) = 0.02798 mol

4. $Cu^{2+} + 2e^- \rightarrow Cu$ needs 2 mol e^- per mol Cu, so n(Cu) = 0.02798 ÷ 2 = 0.01399 mol.

5. m = 0.01399 mol × 63.55 g/mol = 0.889 g

Slide 4 of 8 · Redox titration

A redox titration finds a concentration by reacting a sample with a solution of **known** concentration. Permanganate is reduced: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + \mathbf{4H_2O}$.

Iron(II) is oxidized: $Fe^{2+} \rightarrow Fe^{3+} + e^-$. Five iron(II) ions supply the five electrons that one permanganate needs, so the mole ratio is **1 MnO_4^- to 5 Fe^{2+}**.

Change volumes to **litres** first: divide millilitres by 1000. So 20.00 mL is **0.02000** L and 25.00 mL is 0.02500 L. Then amount = concentration × volume.

The end point is the first **lasting pale pink**. It shows all the iron(II) has reacted. The mole ratio comes from the **coefficients** of the balanced equation, not from 1 to 1.

Example. A 25.00 mL sample of iron(II) solution needs 20.00 mL of 0.0200 mol/L MnO_4^- solution in acid. Find $[Fe^{2+}]$. Given $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and $Fe^{2+} \rightarrow Fe^{3+} + e^-$.

1. Five Fe^{2+} supply $5e^-$: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$

2. Check: Mn 1, O 4, H 8, Fe 5, charge +17 on each side.

3. $n(MnO_4^-) = 0.02000\ L \times 0.0200\ mol/L = 4.00 \times 10^{-4}\ mol$

4. $n(Fe^{2+}) = 5 \times 4.00 \times 10^{-4}\ mol = 2.00 \times 10^{-3}\ mol$

5. $[Fe^{2+}] = 2.00 \times 10^{-3}\ mol \div 0.02500\ L = 0.0800\ mol/L$

Try it. A 20.00 mL sample of iron(II) solution needs 30.00 mL of 0.0100 mol/L MnO_4^- solution in acid. Find $[Fe^{2+}]$. Use $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$.

1. $n(MnO_4^-) = 0.03000\ L \times 0.0100\ mol/L = 3.00 \times 10^{-4}\ mol$

2. $n(Fe^{2+}) = 5 \times 3.00 \times 10^{-4}\ mol = 1.50 \times 10^{-3}\ mol$

3. $[Fe^{2+}] = 1.50 \times 10^{-3}\ mol \div 0.02000\ L = 0.0750\ mol/L$

Slide 5 of 8 · Method: how to choose the method

Step 1. Read what is **given** and what is asked. Do this before you touch a calculator.

Step 2. Sort the problem. A voltage or direction uses the **table**. Current and time use **Faraday**. A reacting volume uses a **titration**.

Step 3. Run that method’s steps. Use the matching method slide from this unit.

Step 4. Check the units and the digits. Round **once**, at the end.

Example. Use the method on a 25.00 mL sample of iron(II) solution that needs 20.00 mL of 0.0200 mol/L MnO_4^- in acid. Find $[Fe^{2+}]$.

1. Step 1: given are two volumes and one concentration. Asked is a concentration.

2. Step 2: a reacting volume points to the titration method.

3. Step 3: $n(MnO_4^-) = 0.02000 \times 0.0200 = 4.00 \times 10^{-4}$ mol. $n(Fe^{2+}) = 5 \times 4.00 \times 10^{-4} = 2.00 \times 10^{-3}$ mol. $[Fe^{2+}] = 2.00 \times 10^{-3} \div 0.02500 = 0.0800$ mol/L.

4. Step 4: the unit is mol/L and the answer has three digits.

Try it. Use the method on 3.00 A for 10.0 minutes through $Ag^+(aq)$ with a silver cathode. M(Ag) = 107.87 g/mol. Choose the method and find the mass of silver.

1. Current and time are given, so use Faraday’s law.

2. $q = 3.00 \times 600 = 1800$ C. $n(e^-) = 1800 \div (9.65 \times 10^4) = 0.018653$ mol. $n(Ag) = 0.018653$ mol.

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

Slide 6 of 8 · Your turn: choose, then solve

Name the method first. Cell A gives table values and asks for a voltage, so use the **E° table**. Cell B gives current and time and asks for a mass, so use **Faraday's law**.

For cell A, zinc has the lower E° (−0.76 V), so zinc is the **anode**. Zinc gives 2 electrons and each silver ion takes 1, so **double** the silver half.

$Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$. $E°cell = 0.80 - (-0.76) =$ **+1.56 V**. This is positive, so the cell is **voltaic**. Do not double the silver value.

For cell B, 30.0 min × 60 = **1800 s**. Sodium ions are reduced at the cathode: $Na^+ + e^- \rightarrow$ **Na**, a ratio of 1 to 1. Write the unit with each answer.

Example. Cell A is a zinc and silver cell: $E°(Zn^{2+}/Zn) = -0.76$ V and $E°(Ag^+/Ag) = +0.80$ V. Cell B is molten NaCl with 2.00 A for 30.0 min. Choose the method for each. Find E°cell for cell A and the mass of sodium for cell B. Use $F = 9.65 \times 10^4$ C/mol e^- and M(Na) = 22.99 g/mol.

1. Cell A asks for a voltage: use the E° table. Cell B gives current and time: use Faraday's law.

2. Cell A: zinc is the anode. $Zn \rightarrow Zn^{2+} + 2e^-$ and $2Ag^+ + 2e^- \rightarrow 2Ag$

3. $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$. $E°cell = 0.80 - (-0.76) = +1.56$ V, so it is voltaic.

4. Cell B: t = 30.0 min × 60 = 1800 s. q = 2.00 A × 1800 s = 3600 C

5. $n(e^-) = 3600 \text{ C} \div (9.65 \times 10^4) = 0.03731$ mol

6. $Na^+ + e^- \rightarrow Na$ is 1 to 1, so m = 0.03731 mol × 22.99 g/mol = 0.858 g

Try it. Cell C is an iron and silver cell: $E°(Fe^{2+}/Fe) = -0.45$ V and $E°(Ag^+/Ag) = +0.80$ V. Cell D is molten NaCl with 2.50 A for 20.0 min. Choose the method for each. Find E°cell for cell C and the mass of sodium for cell D. Use M(Na) = 22.99 g/mol.

1. Cell C asks for a voltage: use the E° table. Cell D gives current and time: use Faraday's law.

2. Cell C: iron is the anode. $Fe + 2Ag^+ \rightarrow Fe^{2+} + 2Ag$

3. $E°cell = 0.80 - (-0.45) = +1.25$ V, so it is voltaic.

4. Cell D: t = 20.0 min × 60 = 1200 s. q = 2.50 A × 1200 s = 3000 C

5. $n(e^-) = 3000 \text{ C} \div (9.65 \times 10^4) = 0.03109$ mol

6. $Na^+ + e^- \rightarrow Na$ is 1 to 1, so m = 0.03109 mol × 22.99 g/mol = 0.715 g

Slide 7 of 8 · Corrosion, industry and batteries

Iron rusts because it is **oxidized**: $Fe \rightarrow Fe^{2+} + 2e^{-}$. A metal with a lower reduction value, like magnesium (−2.37 V), is a stronger **reducing agent** than iron (−0.45 V).

Join magnesium to iron. The magnesium is the **anode** and is used up, so it is a sacrificial anode. The iron is protected as the **cathode**.

In electroplating, the object to plate is the **cathode**. In refining, the impure metal is the **anode**, so it dissolves, and pure metal collects at the cathode.

Every process follows one rule: **oxidation** at the anode and reduction at the **cathode**. A good battery balances voltage, mass, cost, lifetime and disposal.

Example. A steel pipe in soil must be protected from rust. Given $E°(Mg^{2+}/Mg) = -2.37$ V and $E°(Fe^{2+}/Fe) = -0.45$ V, explain how a magnesium block protects the pipe. Name which metal is oxidized and which is the cathode.

1. Magnesium has the lower value (−2.37 V), so it is the stronger reducing agent and is oxidized first.

2. $Mg + Fe^{2+} \rightarrow Mg^{2+} + Fe$ has $E°_{reaction} = -0.45 - (-2.37) = +1.92$ V, which is positive.

3. The magnesium is the sacrificial anode and is used up. The iron pipe is the cathode and is not oxidized.

Try it. Impure copper is refined by electrolysis. Name the electrode that is the impure copper. Write the half-reaction at each electrode.

1. The impure copper is the anode, because oxidation happens there: $Cu \rightarrow Cu^{2+} + 2e^{-}$

2. Pure copper collects at the cathode, because reduction happens there: $Cu^{2+} + 2e^{-} \rightarrow Cu$

Slide 8 of 8 · Redox in industry: bleaching and preserving

Industry bleaches with an **oxidizing** agent. Pulp and paper mills use chlorine dioxide or hydrogen peroxide. Textile makers use hydrogen peroxide or hypochlorite.

The oxidizing agent oxidizes the coloured compounds, so they become **colourless**.

A vat dye such as indigo is treated with a **reducing** agent so that it dissolves. Oxygen in the air then **oxidizes** it back to its colour.

Oxygen is an oxidizing agent. It browns cut fruit and makes fats rancid. An antioxidant such as vitamin C is a **reducing** agent. It reacts with the **oxygen** first, so the food is protected.

Example. A pulp mill bleaches brown wood pulp with hydrogen peroxide. Name the oxidizing agent. Say what it changed and what was oxidized.

1. Hydrogen peroxide is the oxidizing agent. It takes electrons, so it is reduced.

2. The coloured compounds in the pulp lose electrons, so they are oxidized.

3. The oxidized compounds are colourless, so the pulp turns white.

Try it. Vitamin C is added to cut apple slices so they do not turn brown. Is vitamin C an oxidizing agent or a reducing agent? Explain how it protects the apple.

1. Vitamin C is a reducing agent. It gives electrons, so it is oxidized.

2. Oxygen is the oxidizing agent that browns the apple.

3. Vitamin C reacts with the oxygen first, so the apple is not oxidized.

B16 · Worked practice

Unit B review: choose the method | Questions 1-2 of 8

1. Given the Data Booklet table values $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Mg^{2+}/Mg) = -2.37$ V, at 1.0 mol/L. Choose the method. Name the anode and find the voltage.

Use Method: how to choose the method. Use the E° table. Magnesium is the anode. $E°cell = 0.34 - (-2.37) = +2.71$ V.

2. A current of 4.00 A runs for 20.0 min through $Cu^{2+}(aq)$ with a copper cathode. Find the mass of copper that forms. Use M(Cu) = 63.55 g/mol.

$q = 4.00 \times 1200 = 4800$ C. $n(e^-) = 4800 \div (9.65 \times 10^4) = 0.04974$ mol. $n(Cu) = 0.02487$ mol. $m = 1.58$ g.

B16 · Worked practice

Unit B review: choose the method | Questions 3-4 of 8

3. In acid, 20.00 mL of 0.0150 mol/L MnO_4^- reacts with 25.00 mL of an iron(II) solution. Find $[Fe^{2+}]$. Use $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$.

$n(MnO_4^-) = 0.02000 \times 0.0150 = 3.00 \times 10^{-4}$ mol. $n(Fe^{2+}) = 1.50 \times 10^{-3}$ mol. $[Fe^{2+}] = 1.50 \times 10^{-3} \div 0.02500 = 0.0600$ mol/L.

4. A reaction has a calculated E°reaction of −0.50 V. What kind of cell can make it happen, and what does the cell need?

An electrolytic cell. The reaction is not spontaneous, so a power supply must push it with more than 0.50 V.

B16 · Worked practice

Unit B review: choose the method | Questions 5-6 of 8

5. A steel pipe in soil must be protected from rust. Name one method. Say why it works, and say which electrode the pipe is.

Attach a magnesium block. Magnesium has a lower table value than iron, so magnesium is oxidized in place of the iron. The pipe is the cathode.

6. A spoon is to be plated with silver from a silver nitrate solution. Which electrode is the spoon? Write the half-reaction there. Which terminal of the power supply is it joined to?

The spoon is the cathode. $Ag^{+} + e^{-} \rightarrow Ag$. Join the spoon to the negative terminal of the power supply.

B16 · Worked practice

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

7. A mill bleaches wood pulp with hydrogen peroxide. What kind of agent is hydrogen peroxide here? What does it do to the coloured compounds?

Hydrogen peroxide is the oxidizing agent. It oxidizes the coloured compounds to colourless ones, so the pulp turns white.

8. Challenge: Given a zinc and copper voltaic cell that delivers 0.300 A for 2.00 hours. Data Booklet values: $E°(Zn^{2+}/Zn) = -0.76$ V, $E°(Cu^{2+}/Cu) = +0.34$ V and $M(Zn) = 65.41$ g/mol. This question needs two methods. Find $E°_{cell}$ in volts. Then find the mass of zinc that the anode loses, in grams. Use $F = 9.65 \times 10^4$ C/mol e^-.

Voltage needs the E° table. Zinc has the lower value, so it is the anode: $E°_{cell} = +0.34 - (-0.76) = +1.10$ V. Mass needs Faraday's law. Time: $t = 2.00\ h \times 3600\ s/h = 7200$ s. Charge: $q = It = 0.300 \times 7200 = 2160$ C. Moles of electrons: $n(e^-) = 2160 \div (9.65 \times 10^4) = 0.022383$ mol. Anode half-reaction: $Zn \rightarrow Zn^{2+} + 2e^-$, so $n(Zn) = 0.022383 \div 2 = 0.011192$ mol. Mass: $m = 0.011192\ mol \times 65.41\ g/mol = 0.732$ g. The voltage does not enter the mass calculation.