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

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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.

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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.

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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

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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

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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

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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.