EC-010 · MODULE 3 OF 7

Oxidation & Reduction

The matched pair behind every corroding pipe - oxidation and reduction, always running together. The third module of the Electrochemistry & the Galvanic Series set.

EXTERNAL CORROSION TRACK   FOUNDATION   ELECTROCHEMISTRY & THE GALVANIC SERIES   ~30 MINUTES

BY THE END OF THIS MODULE

You'll be able to -

- Define oxidation (loss of electrons) and reduction (gain of electrons) - and explain why oxidation does not require oxygen.
- Explain why the two half-reactions must run together, at matched rates, in electron balance.
- Write the four anodic reactions (Fe, Zn, Mg, Al) and the three cathodic reactions on sight.
- Combine two half-reactions into one balanced overall reaction.
- Tell which cathodic reaction runs in a given environment, and read the field sign that goes with it.
- Explain cathodic protection as the inversion - supplying electrons from outside so iron stops dissolving.

BACKFILLING A MOWER-STRUCK TEST STATION

A magnesium anode, mostly gone. The pipe beside it, clean.

Following the old anode's lead wire down the hole, you find it: a 17-pound magnesium anode that went in as a two-foot bar, now a gnawed gray-white lump, eaten on every face.

A few feet over, the steel it was wired to comes out of the trench clean. Same soil, same water. That anode was put there to disappear so the pipe wouldn't have to - and the question is what that actually means, reaction by reaction.

Oxidation is the metal leaving. Reduction is what takes the electrons it leaves behind. They are locked together - you genuinely cannot have one without the other.

This module is that matched pair.

FIVE SECTIONS, IN ORDER

How we'll work the pair

01
What the words mean
Oxidation, reduction, OIL RIG.

02
Two halves, one process
Simultaneity and electron balance.

03
The oxidation side
The four anodic reactions.

04
The reduction side
The three cathodic reactions.

05
Balance & the inversion
Electron accounting and CP.

SECTION 1 · WHAT THE TWO WORDS MEAN

Loss and gain of electrons

WHERE THE ELECTRONS GO

Oxidation and reduction

Oxidation

The **loss** of electrons. A metal atom lets go of electrons and leaves as a positive ion. The metal-loss side - the corrosion you can see.

$Fe \rightarrow Fe^{2+} + 2e^-$

Reduction

The **gain** of electrons. Something in the environment picks them up. No metal lost - the consuming side that keeps it all running.

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

OIL RIG - Oxidation Is Loss, Reduction Is Gain (of electrons).

IT CATCHES GOOD TECHS

Oxidation does not require oxygen

The name is a historical accident - early chemists noticed metals combining with oxygen. The definition has nothing to do with oxygen being present. Oxidation is the loss of electrons, full stop.

> Steel corrodes in completely oxygen-free ground - waterlogged clay, sealed crevices, the bottom of a flooded trench. The iron still loses electrons; it still oxidizes. What changes without oxygen is the **cathodic** reaction, not the metal loss.

SECTION 2 · TWO HALVES, ONE PROCESS

Locked together by electrons

THE HEART OF THE MODULE

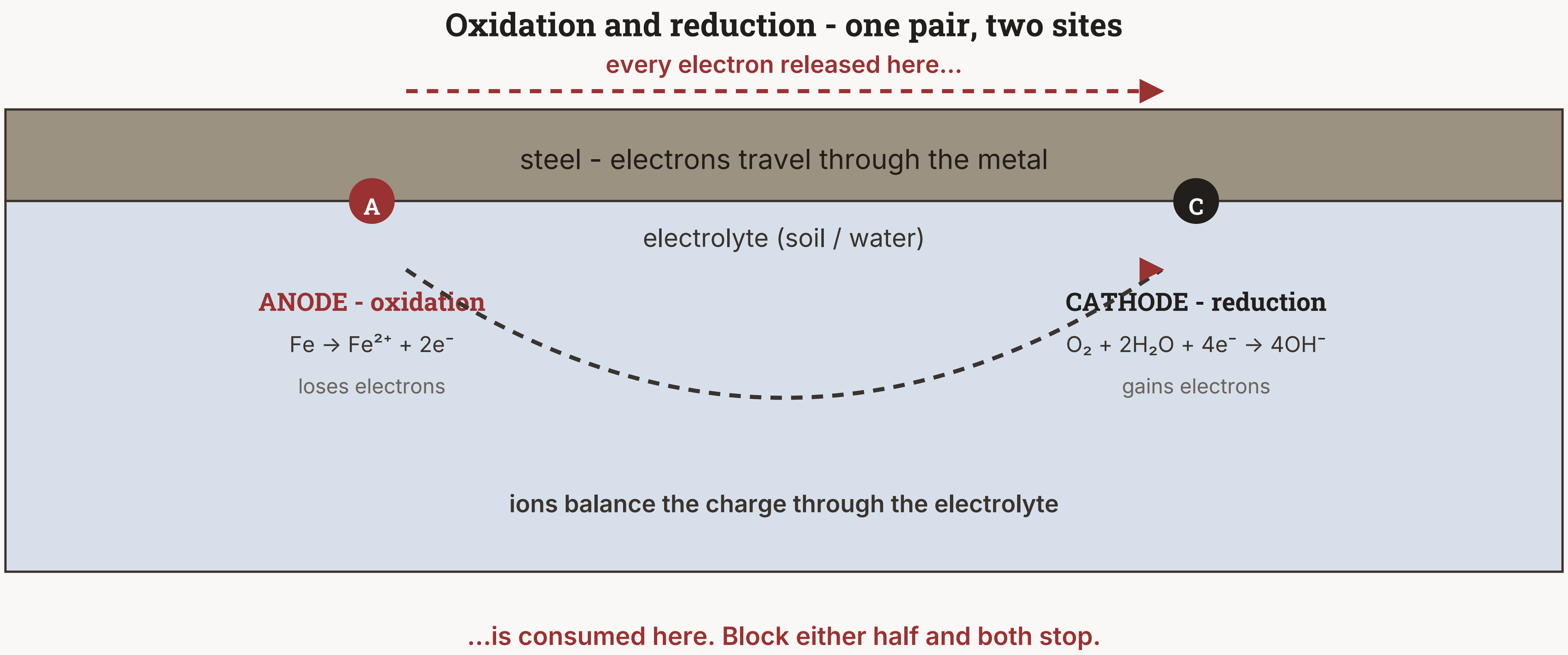
You cannot have one without the other

The two reactions run at the same time, usually at different spots on the metal, at matched rates. Every electron released by oxidation is consumed by a reduction somewhere.

> Block either half and both stop. Speed one up and the other follows. That chain is what every corrosion-control method reaches in to break.

ONE PAIR, TWO SITES

Oxidation here, reduction there



WHY WE WRITE HALVES

The loose electron is the tell

Each reaction carries a free electron - the e^- - so it can't stand alone in the real world. That's why we call each one a half-reaction.

> Two halves add together - one releasing electrons, one consuming them - and the loose electrons cancel, leaving a balanced whole reaction with none left over. When you see e^- in an equation, read it as: **one half of a pair, partner close by.**

A corrosion cell is the same redox chemistry as the battery in your truck. **The only difference is geography** - the two halves are smeared across the same pipe instead of packaged apart.

Same physics. Different distance.

SECTION 3 · THE OXIDATION SIDE

The anodic reactions

EVERY ANODIC REACTION, ONE SHAPE

Metal to ion, electrons left behind

$$M \rightarrow M^{n+} + ne^{-}$$

a metal atom gives up n electrons, becomes an ion, and leaves the surface

The number of electrons, n, is fixed by the metal. This is the only thing that physically removes metal from your pipe - oxidation, every time.

THE ANODIC REACTIONS YOU'LL MEET

One enemy, three friends

Iron	Zinc	Magnesium	Aluminum
$Fe \rightarrow Fe^{2+} + 2e^-$	$Zn \rightarrow Zn^{2+} + 2e^-$	$Mg \rightarrow Mg^{2+} + 2e^-$	$Al \rightarrow Al^{3+} + 3e^-$
YOUR PIPE	ANODE	ANODE	ANODE

Iron is the enemy reaction - your pipe dissolving. The other three are metals we bury on purpose to corrode instead of it. Same move every time: metal to ion, electrons left behind.

SECTION 4 · THE REDUCTION SIDE

The cathodic reactions

THE CATHODE IS A CHAMELEON

Three cathodic reactions

NEUTRAL / AERATED

Oxygen reduction

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

The everyday one for most buried pipe.

ACIDIC / NO OXYGEN

Hydrogen evolution

$2H^+ + 2e^- \rightarrow H_2$

Bubbles off as hydrogen gas.

LESS COMMON

Metal-ion reduction

$Fe^{3+} + e^- \rightarrow Fe^{2+}$

Scaled or chemistry-heavy spots.

THE CATHODE ADAPTS TO THE ENVIRONMENT

The anode never changes; the cathode does

Oxygen present, near neutral

Oxygen reduction runs. Invisible - it just raises the pH and can leave a chalky scale on the protected metal.

Acidic, or oxygen shut out

Hydrogen evolution runs. It has a tell: a fine fizz of hydrogen bubbles on the steel - live reduction you can see.

> The anode runs the identical iron-dissolution reaction in both cases. **Which cathodic reaction is running tells you what kind of place you're standing in.**

AGAINST INTUITION

Oxygen is the fast electron consumer

Oxygen rarely attacks your pipe directly. What it does is keep the cathode running fast - and a fast cathode drags the anode along with it, because the two rates are chained.

That's why well-aerated soil is often more corrosive, not less. Take the fast consumer away and the cell bogs down.

SECTION 5 · BALANCE, AND THE INVERSION

Count the electrons, flip the cell

THE BOOKS ALWAYS BALANCE

Electrons out equal electrons in

Iron gives 2 electrons per atom; oxygen reduction takes 4 per molecule. To make them even, two iron atoms must dissolve for every one oxygen molecule reduced. The chemistry forces the ratio.

> Combine the two halves and the electrons cancel: $2Fe + O_2 + 2H_2O \rightarrow 2Fe(OH)_2$ - the first solid corrosion product, on its way to rust.

WHAT CORROSION IS MEASURED IN

Counted in electrons, not grams

Fifty grams of iron lost is about 1.79 moles of electrons - the same electrons as roughly 59 g of zinc, or 22 g of magnesium.

> Equal electrons, unequal mass, because the metals weigh different amounts per atom. Turning that electron count into pounds of steel and amps of current is the **next module**; here we just establish that the electrons balance.

THE PAYOFF

Cathodic protection is the inversion

The iron only keeps oxidizing because it needs to release electrons, and a reduction nearby keeps taking them. So ask the obvious question: what if the electrons came from somewhere else?

Feed a structure electrons from outside and its iron has no reason to give any up. The metal stops leaving. That's why CP works.

THE LUMP FROM THE START OF THE MODULE

The magnesium was the outside source

It oxidized on purpose, pumping its electrons into the steel so the steel could sit as a cathode and keep its metal. The anode disappeared so the pipe didn't have to.

> Every reading at every test station comes back to one question: is the structure still getting enough electrons from outside that its own iron stays put?

THE METAL WANTS TO GO BACK

Rust is iron returning to what it was

Iron in nature is an ore - an oxide. Refining it into clean steel costs energy, and the metal "wants" to go back to its ore state.

> Corrosion is that slide back downhill, acting itself out. We don't stop the wish - with CP, we just keep paying its electron bill from the outside so it never has to cash in your pipe.

THE CATHODE'S SIGNATURE

Read the cathode, read the ground

Bubbles

Hydrogen evolution - acidic or oxygen-dead ground. Often clean, dark pits, because the iron ions stay dissolved and wash away rather than pile up.

Rust and chalky scale

Oxygen reduction - aerated, near-neutral ground. Hydroxide turns the surface alkaline; rust forms where the iron ions meet oxygen and water.

The anode looks the same everywhere - iron leaving. The cathode's signature is what tells you the environment.

THE TWO QUESTIONS FROM THE START

Where did the metal go? What ate the electrons?

Now you can answer both. The iron went into solution at the anode - oxidation. The electrons it left were consumed at the cathode - reduction. One pair, never apart.

Name the two halves and you can read any corroding surface.

WHAT TO WALK AWAY WITH

The pair, in six lines

- **Oxidation is loss of electrons; reduction is gain.** OIL RIG.
- **Oxidation does not require oxygen** - only the cathodic reaction changes when oxygen is gone.
- **The two halves run together**, at matched rates, in electron balance. Block one and both stop.
- **Four anodic reactions** (Fe, Zn, Mg, Al); **three cathodic** (oxygen reduction, hydrogen evolution, metal-ion).
- **The electrons balance** - two iron atoms per oxygen molecule - and corrosion is counted in electrons, not grams.
- **CP is the inversion:** supply electrons from outside and the iron stops dissolving.

WHY THIS PAIR IS THE HINGE

Control the electrons, control the corrosion

Every way we fight corrosion reaches in and breaks the electron chain - shrink the exposed metal, cut the electron path, or feed the structure electrons from outside.

Oxidation and reduction are the move the whole game is played with. You just learned it.

WHERE THE ELECTROCHEMISTRY SET GOES NEXT

Each module is a closer look at this pair

- **EC-011** - the galvanic series - which metal goes anodic to which, and by how much.
- **EC-012** - corrosion rate - turning the electron balance into pounds of metal and amps of current.
- **EC-013** - reference electrodes - the tools that measure the cell's potentials.
- **EC-014** - polarization - how the reaction rates shift once the cell is running.

FURTHER READING + THE NEXT MODULE

Sources used + where this goes

- **The Electrochemical Reaction Leading to Corrosion & Reactions at the Anode Surface** - Field Notes from RCS (this module's supplement)
- **Electrochemical Basics** - Field Notes from RCS
- **Corrosion Basics: An Introduction**
- **Cathodic Protection Training Materials** (CP-1 / CP-2)
- **AUCSC Short Course Materials**
- **Peabody's Control of Pipeline Corrosion**
- **NACE SP0169** - Control of External Corrosion on Underground or Submerged Metallic Piping Systems

UP NEXT

EC-011 - The Galvanic Series

Which metal in a pair becomes the anode, and by how much - putting a ranking to the "more active metal oxidizes" rule you used all module.