ChemBam

Imaging metal corrosion (teacher guide to exercises and experiments)

AIM

- For students to appreciate that metal corrosion occurs all around us, affecting building, pipes, containers for nuclear waste.
- For students to understand that corrosion is a spontaneous electrochemical process that happens in aqueous solutions, and therefore occurs when metals are exposed to water.
- For students to observe the redox reactions that occur at a metal, by visualising the change in pH in agar; blue/purple areas show reduction of oxygen in the air, to produce hydroxide, and orage/red areas show the corrosion of the iron to form acidic Fe^{3+} complexes in solution.
- For students to appreciate that an electrochemical cell sets itself up spontaneously, when metals are exposed to air/water/salt, and that salt speeds up the corrosion by helping the movement of ions in solution.
- Students can appreciate that different metals corrode differently, either from the perspective of reactivity series of metals, or A-Level electrochemistry and standard electrode potentials.
- For students to appreciate how galvanising metals can protect them from corrosion.

EXPERIMENT QUESTIONS– answer guide for teachers

1. This works best as a 'class' experiment, and across the class you can compare the observations of the different combinations of metals, to create a table:

Metal	Electrochemistry	Observation
Copper	Copper is harder to oxidise than iron: E^{o}_{cell} =0.40 - +0.34 = +0.06 V Copper is less reactive than iron	Very little is observed with the copper, it corrodes very slowly.
Iron (steel nail)	Iron oxidises much more readily than copper E^{o}_{cell} =0.40 - -0.44 = +0.84 V Iron is more reactive than copper	The iron reduced the oxygen much faster, and areas of blue/purple (UI) or pink (PI) indicate reduction of oxygen. The nail will start to show orange/brown rust forming.
Iron wrapped in copper	Copper is less reactive than iron, (has a more positive standard electrode potential) and pulls electrons away from iron. This makes iron more likely to oxidise and corrode.	The iron corrodes faster, especially where the copper wire is in contact with the iron.
Iron wrapped in zinc	Zinc is more reactive than iron, (has a more negative standard electrode potential) and gives up its elecrons to the iron. Zinc preferentially gives up electrons to the oxygen, and pushes electrons towards the iron, protecting it from corrosion.	The iron corrodes slower, and reduction is mainly around the zinc wire. Reduction still occurs readily, and OH^- forms, but the nail remains without rust.

2. Fe^{2+} and Fe^{3+} ions form in the agar, and the universal indicator turns red and orange, as the Fe ions form acidic species. Gradually orange/brown solid Fe_2O_3 (rust) forms.
3. Copper is less reactive than iron, (has a more positive standard electrode potential) and pulls electrons away from iron. This makes iron more likely to oxidise and corrode. Zinc is more reactive than iron, (has a more negative standard electrode potential) and gives up its electrons to the iron. Zinc preferentially gives up electrons to the oxygen, and pushes electrons towards the iron, protecting it from corrosion. This method of 'galvanising' is used to protect nails, buckets, boats (etc) from rusting, by coating the steel in a coating of zinc.

4. Salt speeds up the rate of corrosion, because it forms ions in solution, i.e. an electrolyte, which conducts ions, allowing electrons to flow, and hence speeding up the electrochemical process.