PRAC 4 – REDOX CHEMISTRY (WEEK 6)

Zinc reacts with HCl acid giving off H_2 gas

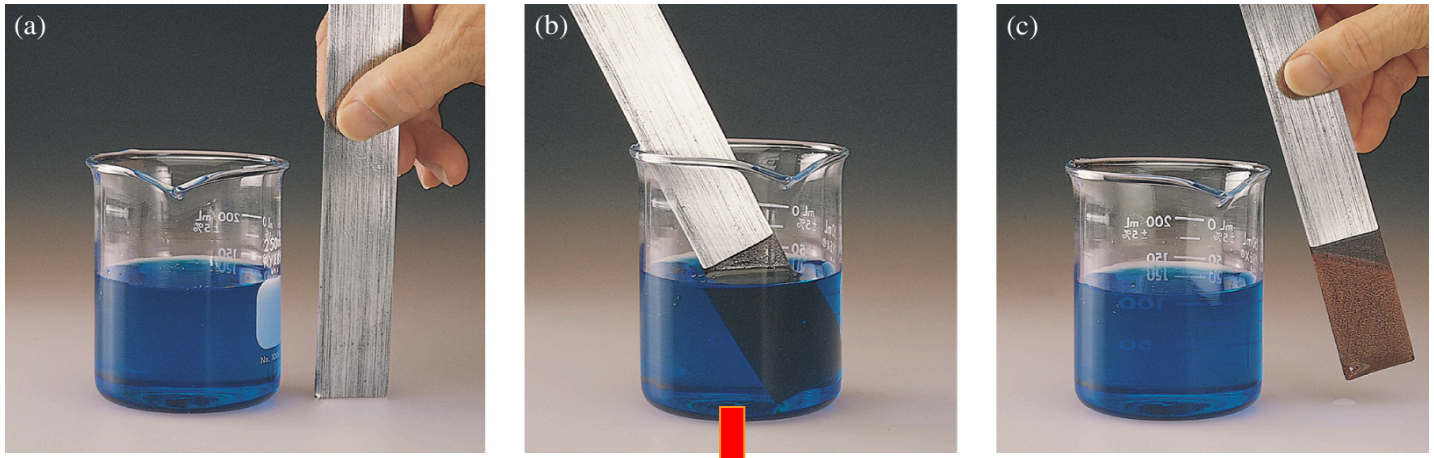
FROM WEEK 6 LECTURES

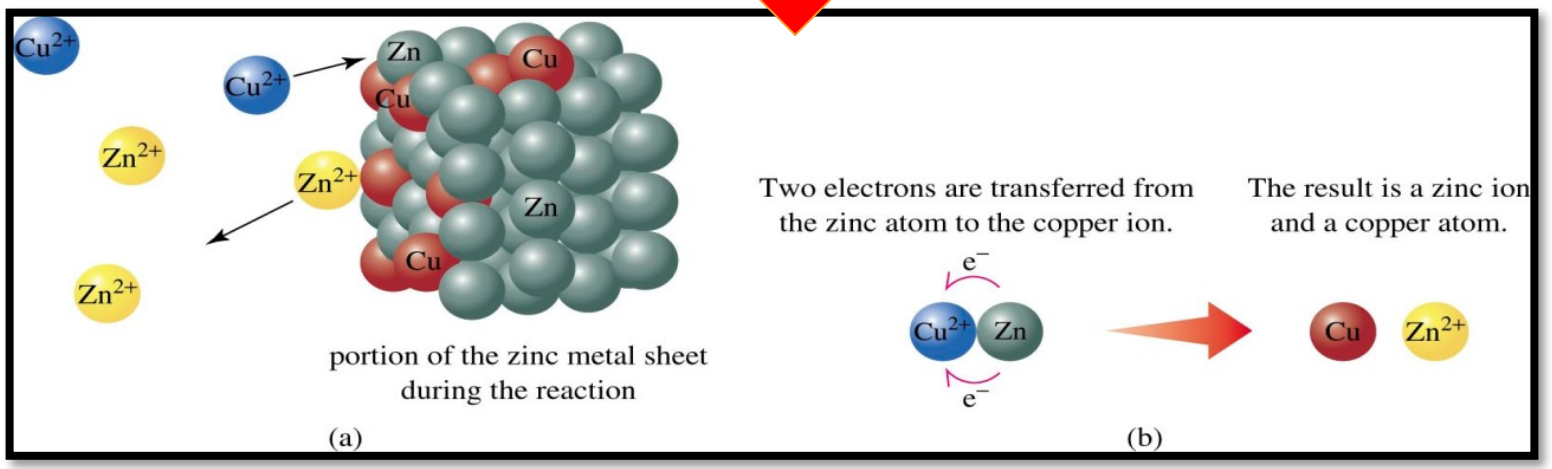
- An oxidation cannot occur without a simultaneous reduction and vise versa
- One substance must release electrons (oxidised)
- The other must accept the electrons (reduced)
- If oxidants and reductants are separated by a conducting wire we can obtain a useful current
- A **galvanic cell** derives electrical energy from a **spontaneous** redox reaction

12.3 GALVANIC CELLS

Chemical Change ⟫ Electric Current

➢ $Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)$

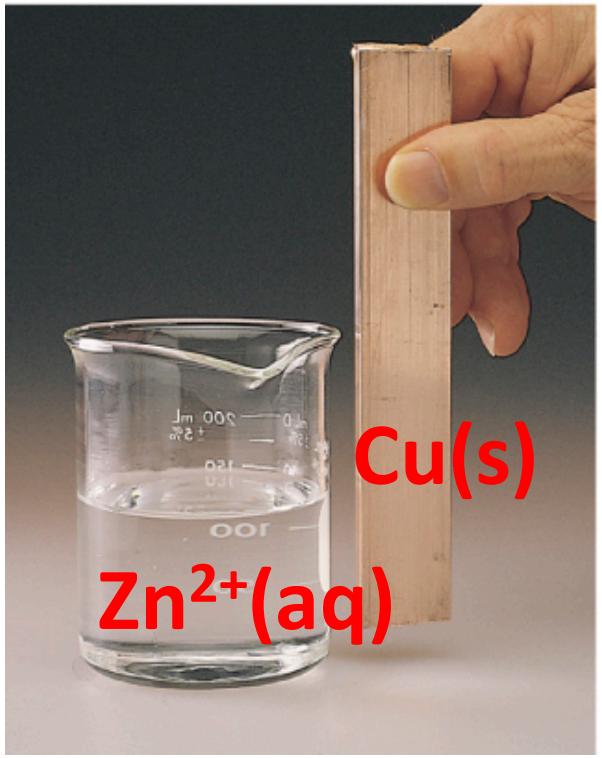




12.3 GALVANIC CELLS

Why??

$$Cu(s) + ZnSO_4(aq) \longrightarrow \text{no reaction}$$



Copper cannot reduce zinc ions to metallic zinc therefor no reaction

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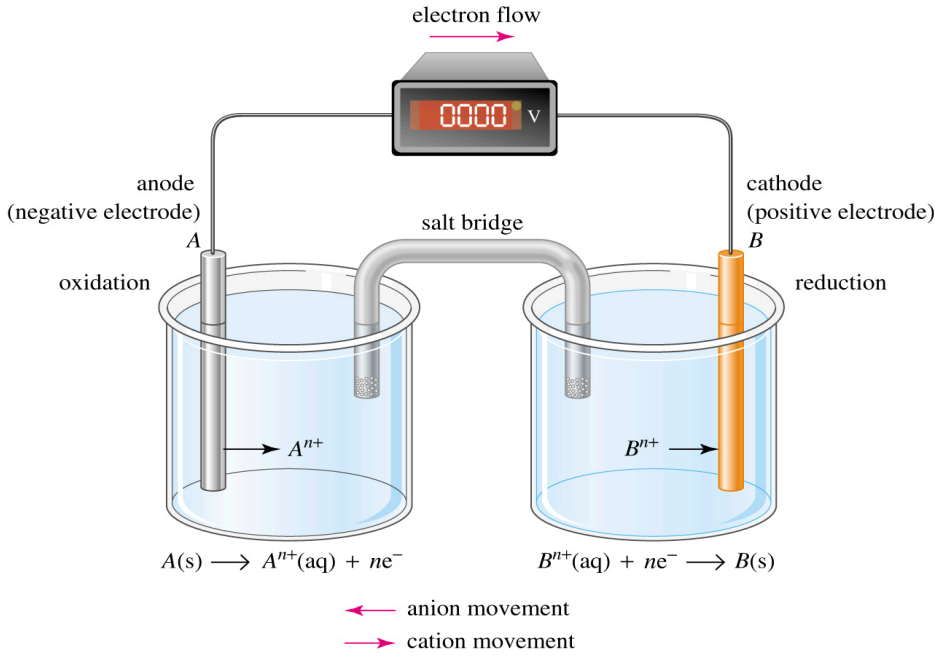
12.4 REDUCTION POTENTIALS

When two half-cells are connected:

- The one with the larger reduction potential will acquire electrons and undergo reduction
- The half-cell with the lower reduction potential will give up electrons and undergo oxidation



In this prac we are not separating reaction in cells.

Part A - corrosion

A solution of 3 % sodium chloride (salt water to facilitate conductance)

0.2 % potassium ferricyanide (turns blue when Fe^{2+} is made = oxidation)

0.1 % phenolphthalein (turns pink when OH^- made = reduction)

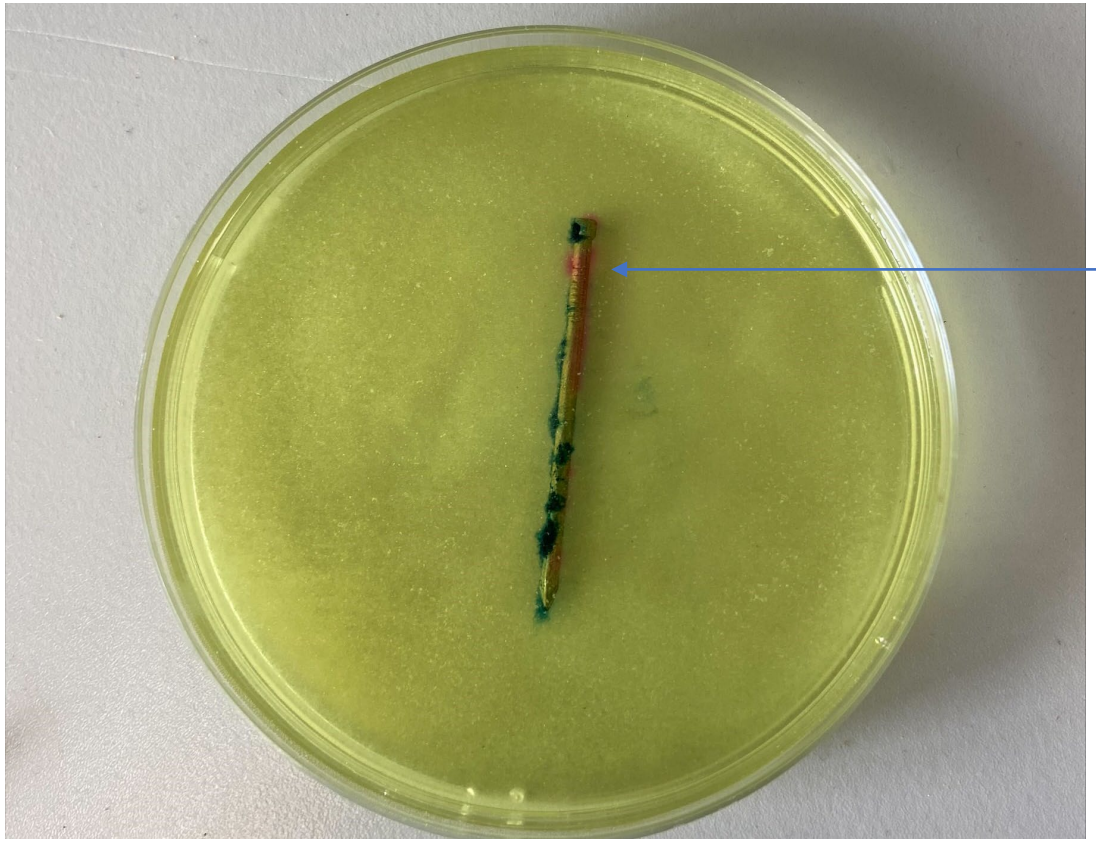
1.0 % agar – creates homogenous jelly like plate with above reagents

Pink colour represents OH^- formed, so another substance has been oxidised
A blue colour forms when the ferricyanide complex is formed, Fe is oxidised to Fe^{2+}

Textbook section 12.5 has more information regarding corrosion

Nail 1: Plain iron (Fe) nail

Fe nail in oxygenated salt water with indicator
Small amount of blue indicating Fe^{2+} formed,
Pink formed in different location, indicating OH^- formed.

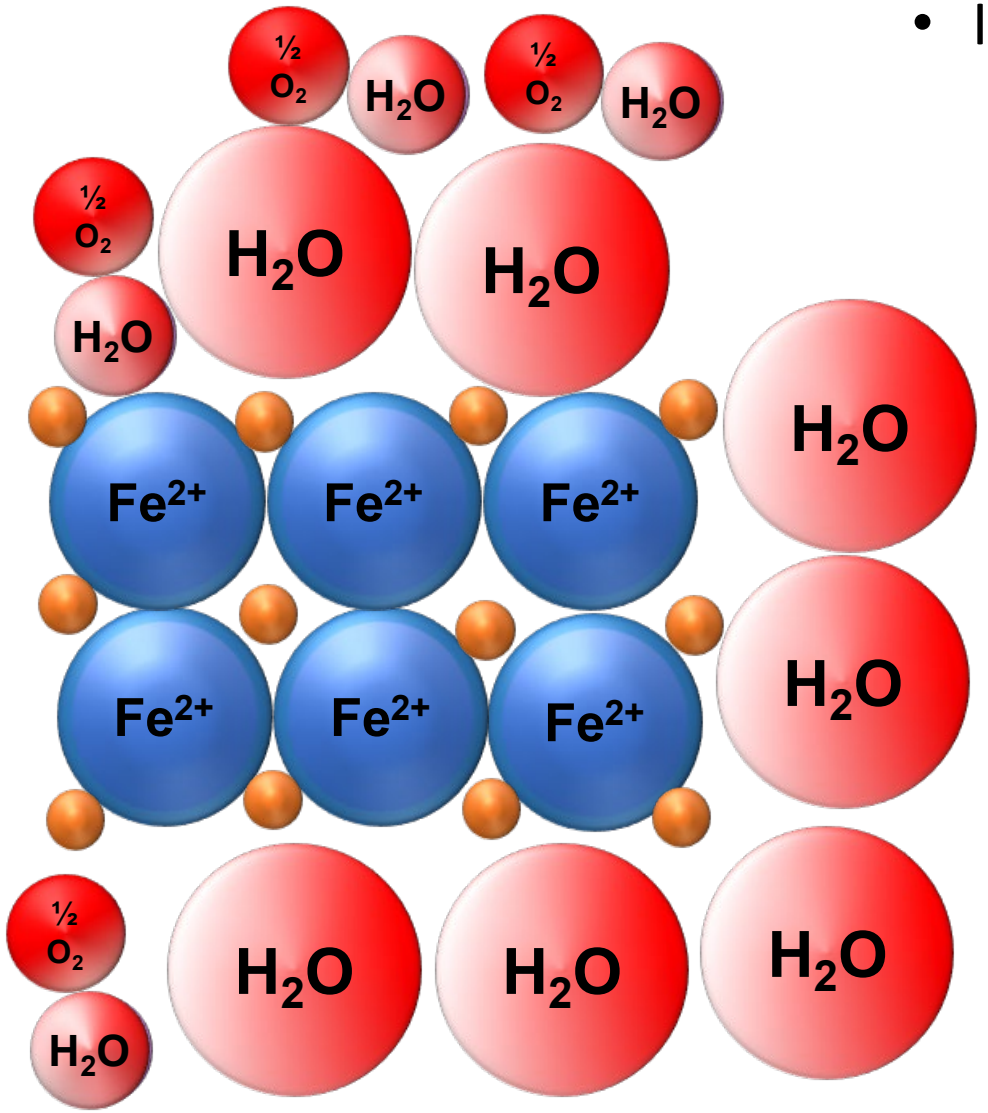


Note pink and blue in different areas

Oxidation and reduction have to take place in different locations

https://edu.rsc.org/exhibition-chemistry/nailing-corrosion-demonstrations/2000054.article

Oxygenated water near surface – able to take electrons from iron



- Iron ions dissolve
- Indicator turns blue

- Oxygenated water accepts electrons
- Hydroxide ions formed
- Indicator turns pink

$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$ -0.45 V

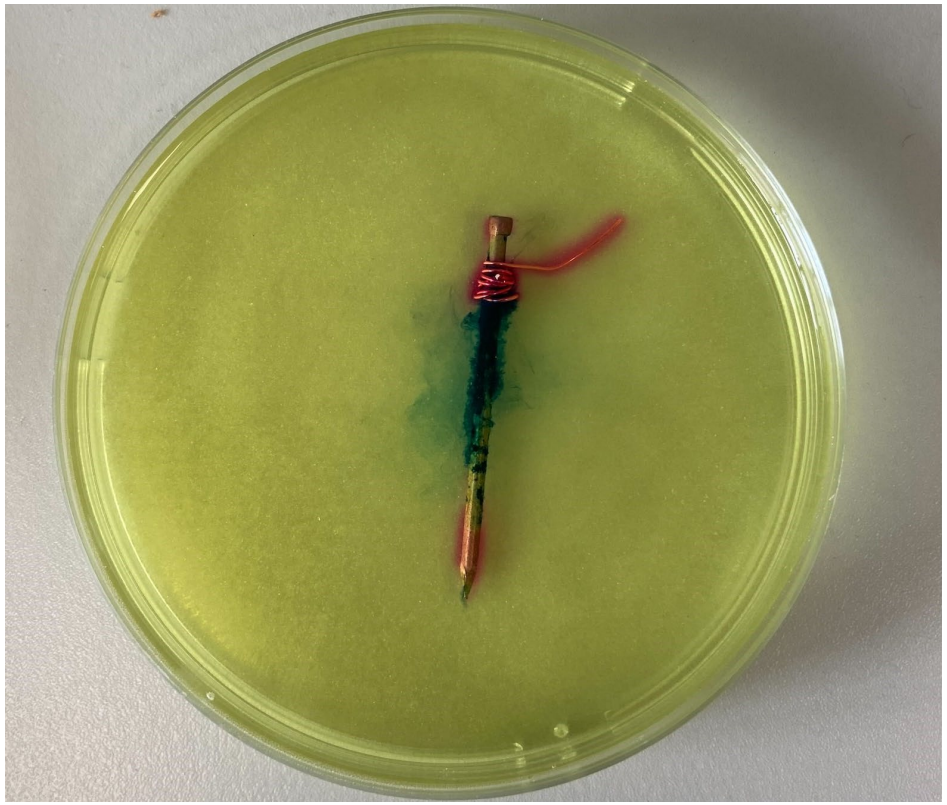
$\frac{1}{2}O_2(g) + H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq)$ +0.40 V

What needs to happen next … ?

https://edu.rsc.org/exhibition-chemistry/nailing-corrosion-demonstrations/2000054.article

Nail 2: Plain Fe nail with Cu wrapped around

Fe nail and Cu wire, Fe^{2+} formed at nail (anode), OH^- formed at cathode (Cu)



Cu is the cathode, electrons transferred to but $O_2 + H_2O$ are reduced

Cu already fully reduced so it simply transferred electrons

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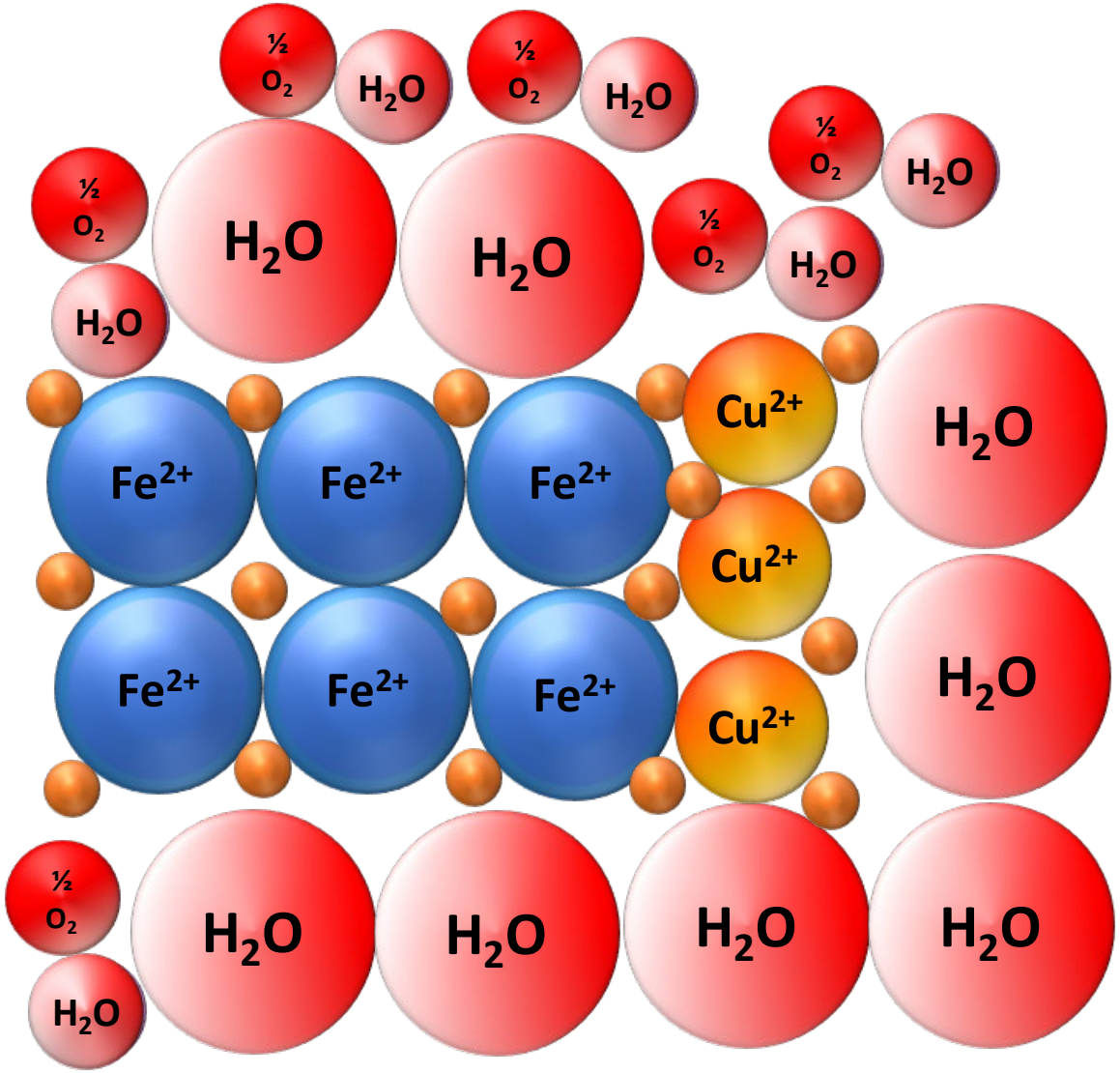
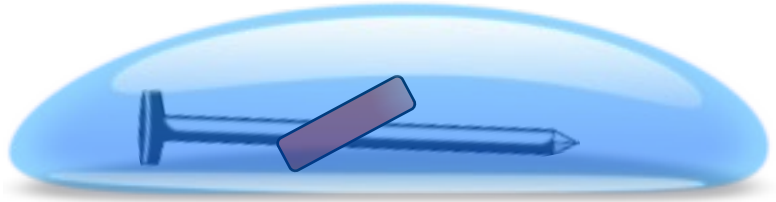
Fe nail wrapped in cupper metal





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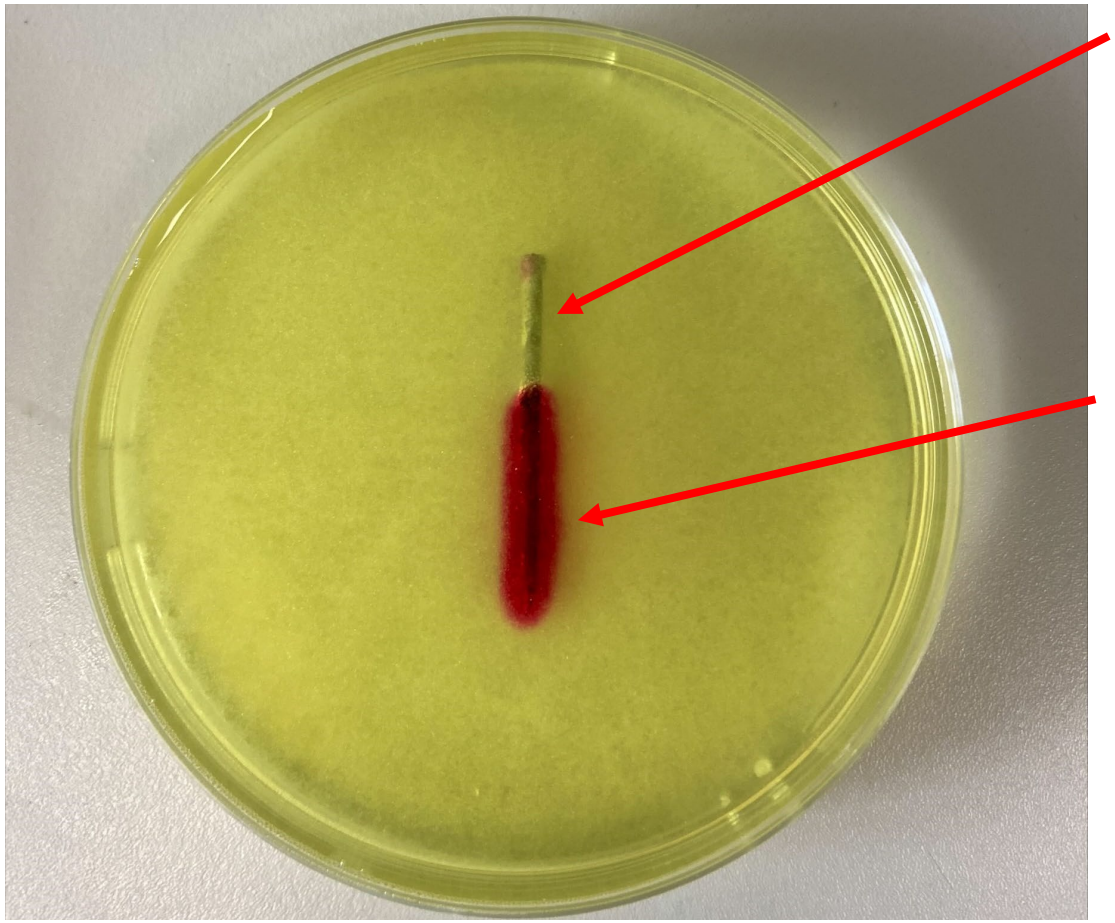
https://edu.rsc.org/exhibition-chemistry/nailing-corrosion-demonstrations/2000054.article

Nail 3: Galvanised nail, Fe coated in Zn

Zn coating removed at the tip of the nail so that **BOTH** Fe and Zn are present

Check out electrochemical series, predict which element will oxidise



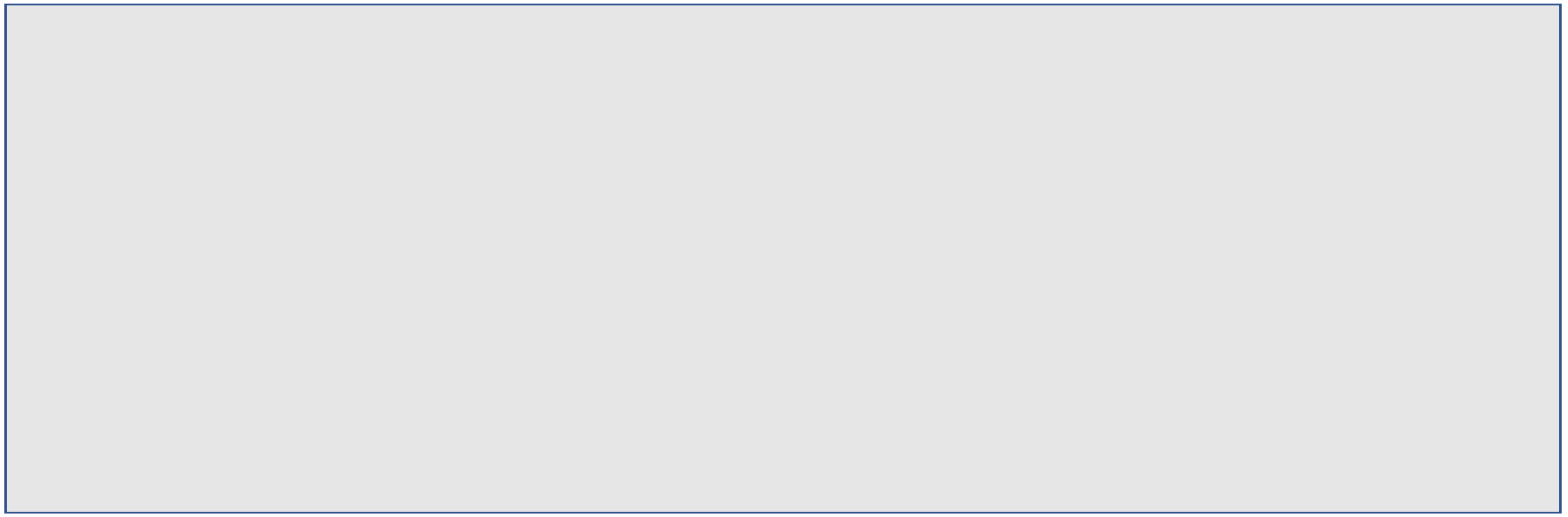
Zn oxidised here, Zn^{2+} doesn't form blue colour

Fe acts as the cathode, and is fully reduced (no blue formed)

OH^- formed where Fe is exposed

Reaction for nail 3, note no reaction for Fe(s) in the nail, it is acting as the cathode and transferring electrons to reduce water and oxygen.

It is the zinc that is getting oxidised, OH^- formed at the tip where the Fe metal is exposed, transferring electrons to H_2O and O_2, creating OH^-.



Part B: reactions of metals with acid

Link this to the electrochemical series, on pg 716 of the textbook

$NiO_2(s) + 2H_2O(l) + 2e^-$	⇌	$Ni(OH)_2(s) + 2OH^-(aq)$	+0.49
$Cu^{2+}(aq) + 2e^-$	⇌	$Cu(s)$	+0.34
$Cl_2(aq) + 4OH^-(aq)$	⇌	$2OCl^-(aq) + 2H_2O(l) + 2e^-$	+0.32
$Hg_2Cl_2(s) + 2e^-$	⇌	$2Hg(l) + 2Cl^-(aq)$	+0.27
$AgCl(s) + e^-$	⇌	$Ag(s) + Cl^-(aq)$	+0.23
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$	⇌	$H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^-$	⇌	$Sn^{2+}(aq)$	+0.15
$AgBr(s) + e^-$	⇌	$Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^-$	⇌	$H_2(g)$	0
$Sn^{2+}(aq) + 2e^-$	⇌	$Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^-$	⇌	$Ni(s)$	−0.25
$Co^{2+}(aq) + 2e^-$	⇌	$Co(s)$	−0.28
$PbSO_4(s) + H^+(aq) + 2e^-$	⇌	$Pb(s) + HSO_4^-(aq)$	−0.36
$Cd^{2+}(aq) + 2e^-$	⇌	$Cd(s)$	−0.40
$Fe^{2+}(aq) + 2e^-$	⇌	$Fe(s)$	−0.44
$Cr^{3+}(aq) + 3e^-$	⇌	$Cr(s)$	−0.74
$Zn^{2+}(aq) + 2e^-$	⇌	$Zn(s)$	−0.76
$2H_2O(l) + 2e^-$	⇌	$H_2(g) + 2OH^-(aq)$	−0.83
$Al^{3+}(aq) + 3e^-$	⇌	$Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^-$	⇌	$Mg(s)$	−2.37
$Na^+(aq) + e^-$	⇌	$Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^-$	⇌	$Ca(s)$	−2.76
$K^+(aq) + e^-$	⇌	$K(s)$	−2.92
$Li^+(aq) + e^-$	⇌	$Li(s)$	−3.05

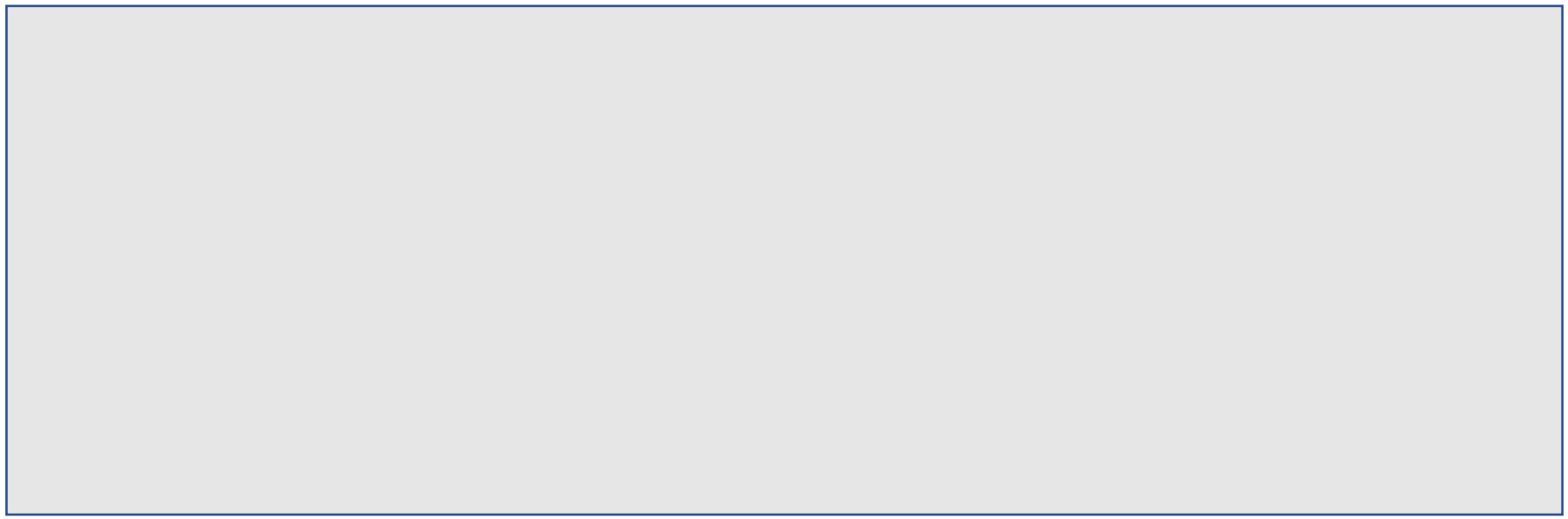
Weakest oxidant

Strongest reductant

Metals higher will not react with the acid (Cu)

Metals below will react.
Note Fe and Sn hard to see in the video

Use the electrochemical series to help write the net ionic equations



Part C: Reactions of metals with solutions of metal ions

- Some reactions hard to see, use the electrochemical series to help
- Al has oxide coating so does not react as predicted, table has been filled in for Al
- If solution containing metal ion (eg Cu^{2+}) is higher on the series then it should react with the solid metal