

Lecture Note #16 (Spring, 2022)

Corrosion

1. Corrosion fundamentals
2. Thermodynamics of corrosion
3. Corrosion rate
4. Localized corrosion
5. Corrosion protection

Corrosion fundamentals

- Corrosion: **unwanted attack of metals by their environment**

e.g. $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$

- **Components** of corrosion systems

(1) The metal (anode)

(2) A cathode reactant with an equilibrium potential higher than that of the metal

(3) Electrical contact between the anode and cathode

(4) An electrolyte that permits current flow in solution between the anode and cathode

e.g.

galvanic corrosion (like Daniel cell (battery)): an active metal is electrically connected to a more noble metal (one that is less likely to corrode) → oxidation in one side + reduction in the other side

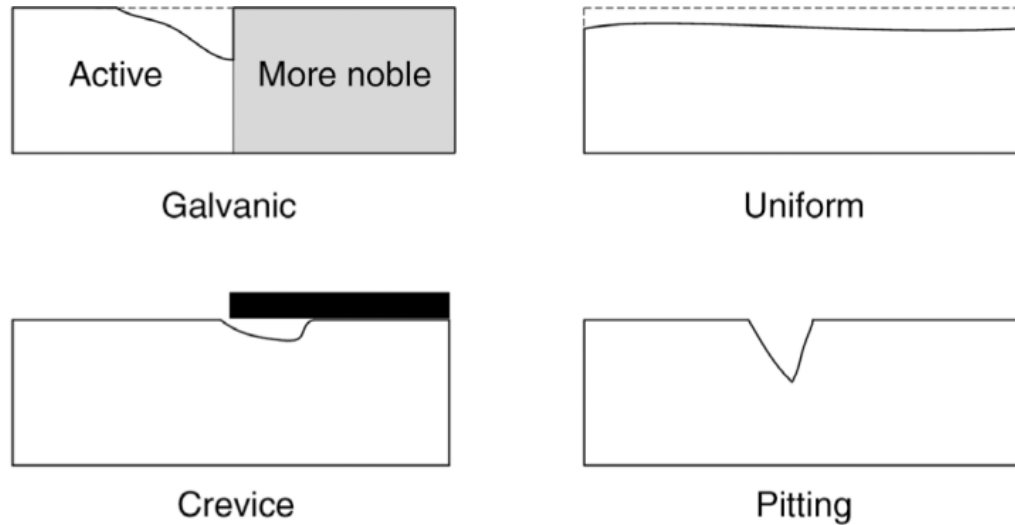
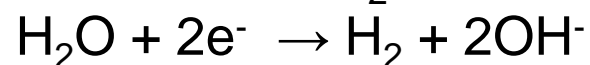
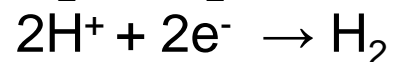
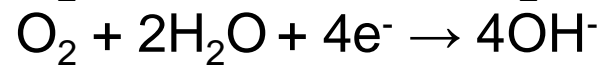
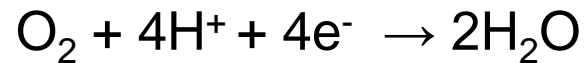


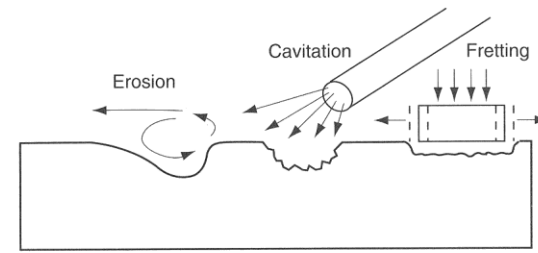
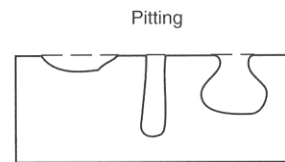
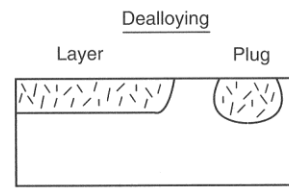
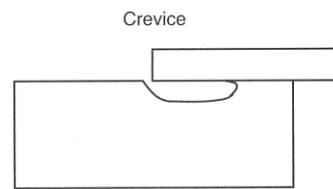
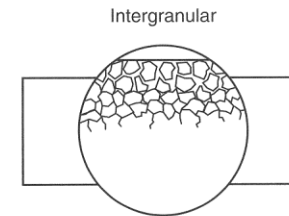
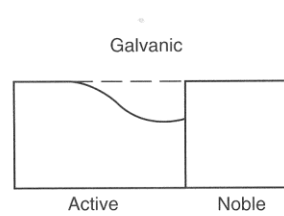
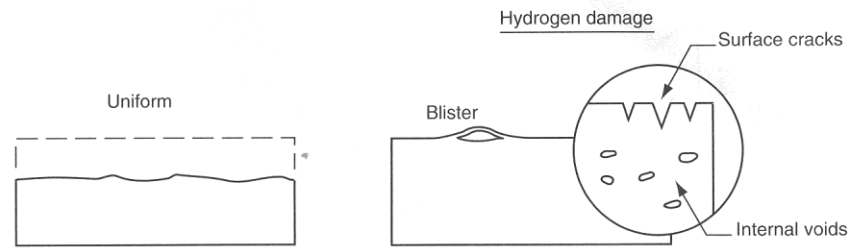
Figure 16.1 Examples of different types of corrosion.

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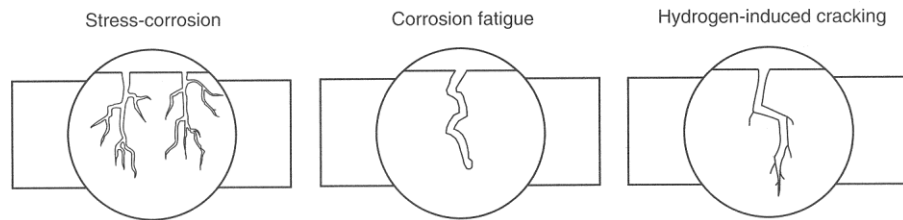
- Most common cathodic reactions in corrosion in aqueous systems



Types of corrosion



Environmentally Induced Cracking



Thermodynamics of corrosion systems

- Thermodynamically favorable corrosion reaction is a reaction that has a **negative free energy (positive cell potential)** with the metal as the anode
- **Immunity**: corrosion is not thermodynamically favorable
- **Passive layer**: Al is very reactive (standard potential = -1.66 V), but doesn't corrode. Why? Oxide layer → passive layer, **passivation**
- potential, pH: critical variable in determining the stability of the surface layer → **Porbaix diagram**

Standard electrode potential (표준전극전위)

E°

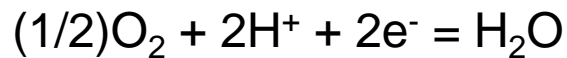
	Reaction	Standard Potential, e° (volts vs. SHE)
Noble	$\text{Au}^{3+} + 3e^{-} = \text{Au}$	+1.498
	$\text{Cl}_2 + 2e^{-} = 2\text{Cl}^{-}$	+1.358
	$\text{O}_2 + 4\text{H}^{+} + 4e^{-} = 2\text{H}_2\text{O}$ (pH 0)	+1.229
	$\text{Pt}^{3+} + 3e^{-} = \text{Pt}$	+1.2
	$\text{O}_2 + 2\text{H}_2\text{O} + 4e^{-} = 4\text{OH}^{-}$ (pH 7) ^a	+0.82
	$\text{Ag}^{+} + e^{-} = \text{Ag}$	+0.799
	$\text{Hg}_2^{2+} + 2e^{-} = 2\text{Hg}$	+0.788
	$\text{Fe}^{3+} + e^{-} = \text{Fe}^{2+}$	+0.771
	$\text{O}_2 + 2\text{H}_2\text{O} + 4e^{-} = 4\text{OH}^{-}$ (pH 14)	+0.401
	$\text{Cu}^{2+} + 2e^{-} = \text{Cu}$	+0.337
	$\text{Sn}^{4+} + 2e^{-} = \text{Sn}^{2+}$	+0.15
	$2\text{H}^{+} + 2e^{-} = \text{H}_2$	0.000
	$\text{Pb}^{2+} + 2e^{-} = \text{Pb}$	-0.126
	$\text{Sn}^{2+} + 2e^{-} = \text{Sn}$	-0.136
	$\text{Ni}^{2+} + 2e^{-} = \text{Ni}$	-0.250
	$\text{Co}^{2+} + 2e^{-} = \text{Co}$	-0.277
	$\text{Cd}^{2+} + 2e^{-} = \text{Cd}$	-0.403
	$\text{Fe}^{2+} + 2e^{-} = \text{Fe}$	-0.440
	$\text{Cr}^{3+} + 3e^{-} = \text{Cr}$	-0.744
$\text{Zn}^{2+} + 2e^{-} = \text{Zn}$	-0.763	
$2\text{H}_2\text{O} + 2e^{-} = \text{H}_2 + 2\text{OH}^{-}$	-0.828	
$\text{Al}^{3+} + 3e^{-} = \text{Al}$	-1.662	
$\text{Mg}^{2+} + 2e^{-} = \text{Mg}$	-2.363	
$\text{Na}^{+} + e^{-} = \text{Na}$	-2.714	
Active	$\text{K}^{+} + e^{-} = \text{K}$	-2.925

Pourbaix diagrams

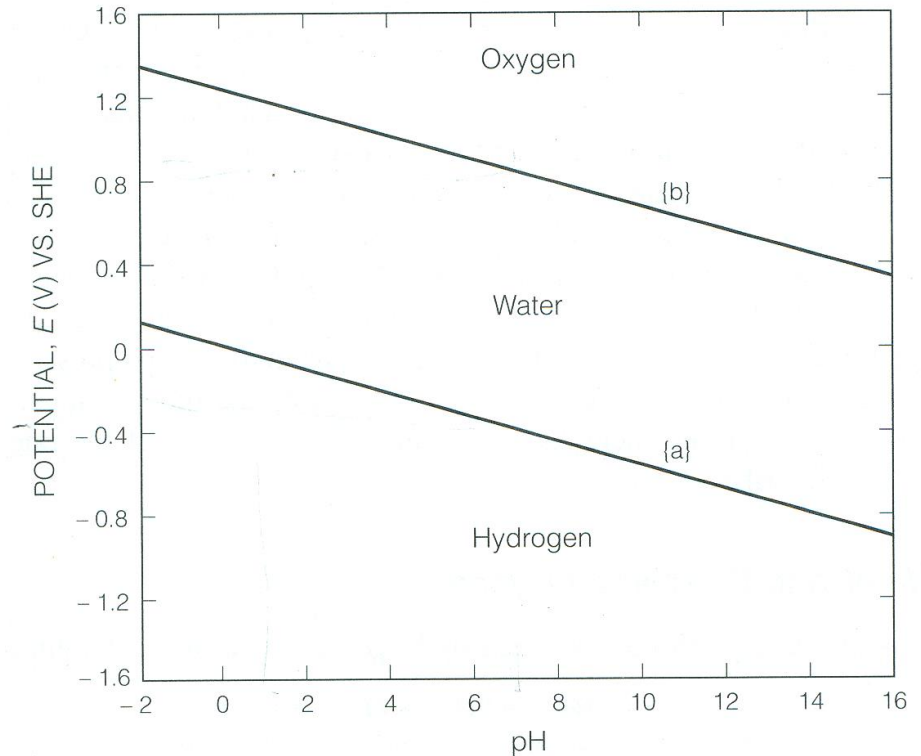
at 25°C, $\text{pH} = -\log_{10} a$



$$\begin{aligned} E &= E^0 - (RT/2F)\ln(1/a_{\text{H}^+}^2) \\ &= 0 + (RT/F)\ln(a_{\text{H}^+}) \\ &= -(RT/F)(2.303\text{pH}) \\ &= -0.0592\text{pH} \end{aligned}$$



$$\begin{aligned} E &= E^0 - (RT/2F)\ln(1/a_{\text{H}^+}^2) \\ &= 1.229 + (RT/F)\ln(a_{\text{H}^+}) \\ &= 1.229 - (RT/F)(2.303\text{pH}) \\ &= 1.229 - 0.0592\text{pH} \end{aligned}$$

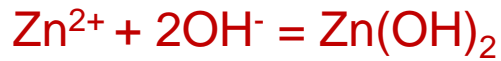


Pourbaix diagram for Zn at 25°C



$$\begin{aligned} E &= E^0 - (RT/2F)\ln(1/a_{\text{Zn}^{2+}}) \\ &= -0.763 + (RT/2F)\ln(a_{\text{Zn}^{2+}}) \end{aligned}$$

$$[\text{Zn}^{2+}] = 10^{-6} \text{ M} \rightarrow -0.94 \text{ V (line c)}$$

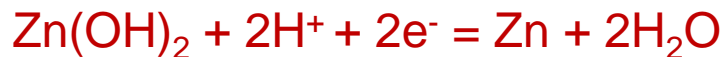


No electron transfer reaction (not a function of potential)

→ vertical line (line d)

→ pH calculation using K_{sp}

→ pH = 8,74



$$\begin{aligned} E &= -0.425 + (RT/F)(2.303\text{pH}) \\ &\text{(line e)} \end{aligned}$$

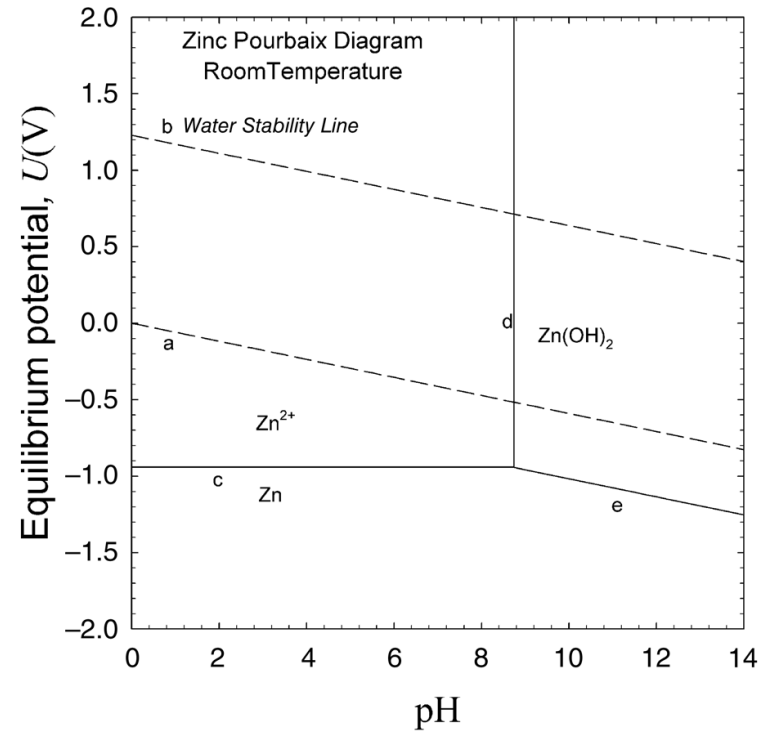


Figure 2.2 Simplified Pourbaix diagram for Zn.

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Ni

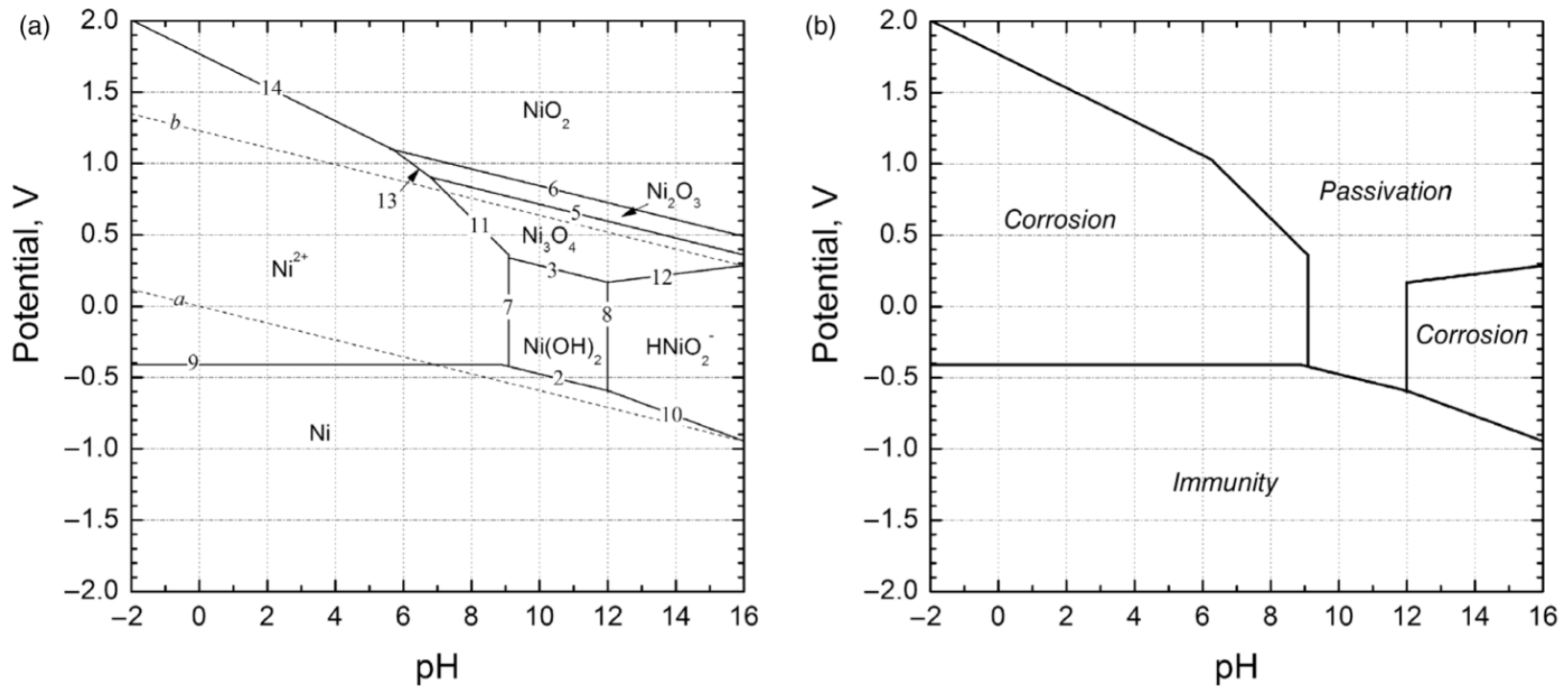
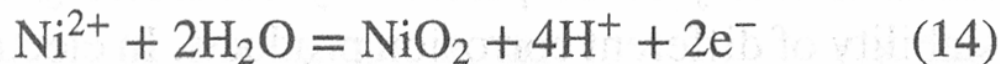
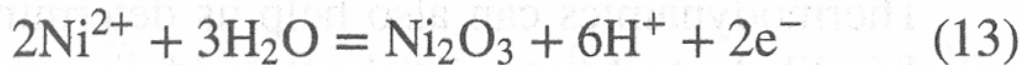
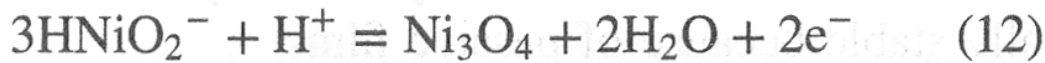
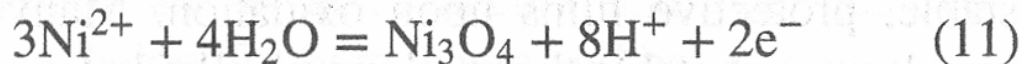
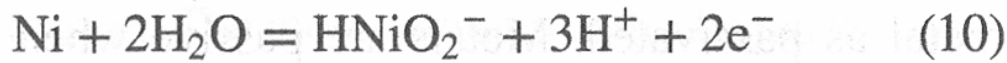
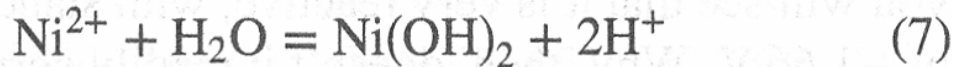
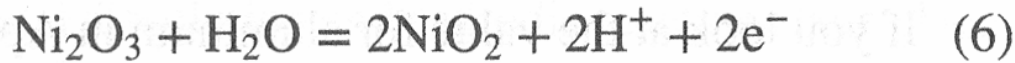
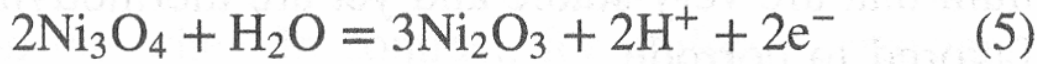
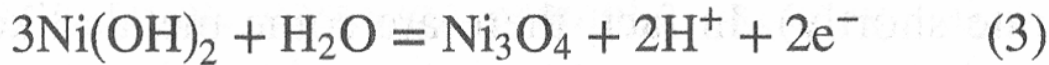


Figure 16.2 Pourbaix diagram for nickel and resulting corrosion map showing regions of immunity, corrosion, and passivation. The numbers and letters correspond to the equations below and in those used in Pourbaix's book.

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- Dashed line a , b: hydrogen reaction and oxygen reaction
- Each solid line: equilibrium of an electrochemical or a chemical reaction



From ch.13 Electrodeposition

- Anode is made from the same metal as that being plated at the cathode and condition is selected that the concentration of the metal ion in solution remains constant
- Anode and cathode current efficiencies to be matched
- From Fig. 8.11, in a plating process, it is essential to maintain the anode potential in the active region to ensure smooth and uniform dissolution (maintaining a suitable low current density)
- Additions of complexing agents (preventing passivation) or halide ions (causing pitting of the oxide film)

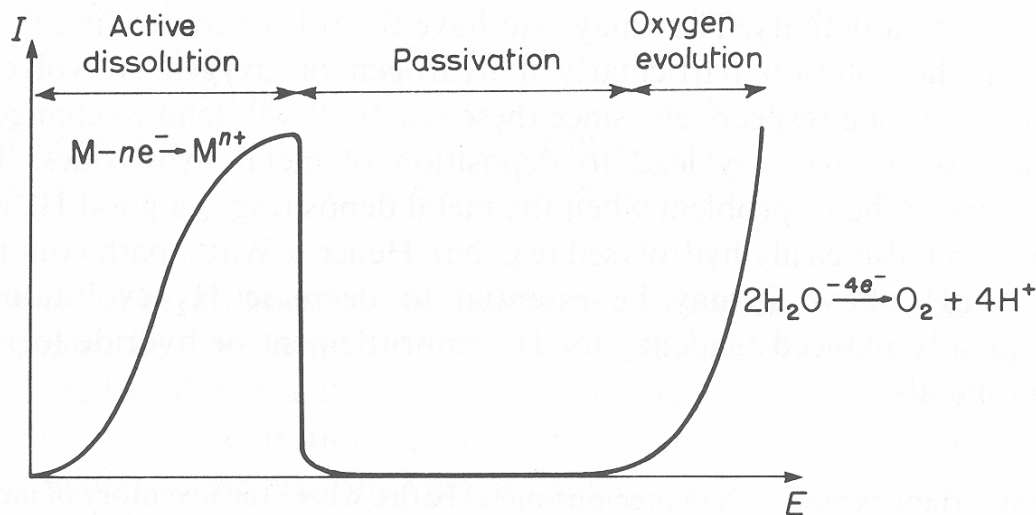


Fig. 8.11 Typical current-potential curve for an anode in an electroplating bath.

Some metals (eg Cr) are passivated when used as an anode and cannot be eluted. Metal ions are generated by dissolving the salt of the metal to be electrodeposited in the electrolyte.

일부 금속(예, Cr)은 산화전극(anode)으로 사용할 경우 부동화(passivation)되어 용출이 불가능. 전착하고자 하는 금속의 염을 전해액에 용해하여 금속 이온 생성시킴

- Vertical line in Pourbaix diagram (e.g. pH = 8)

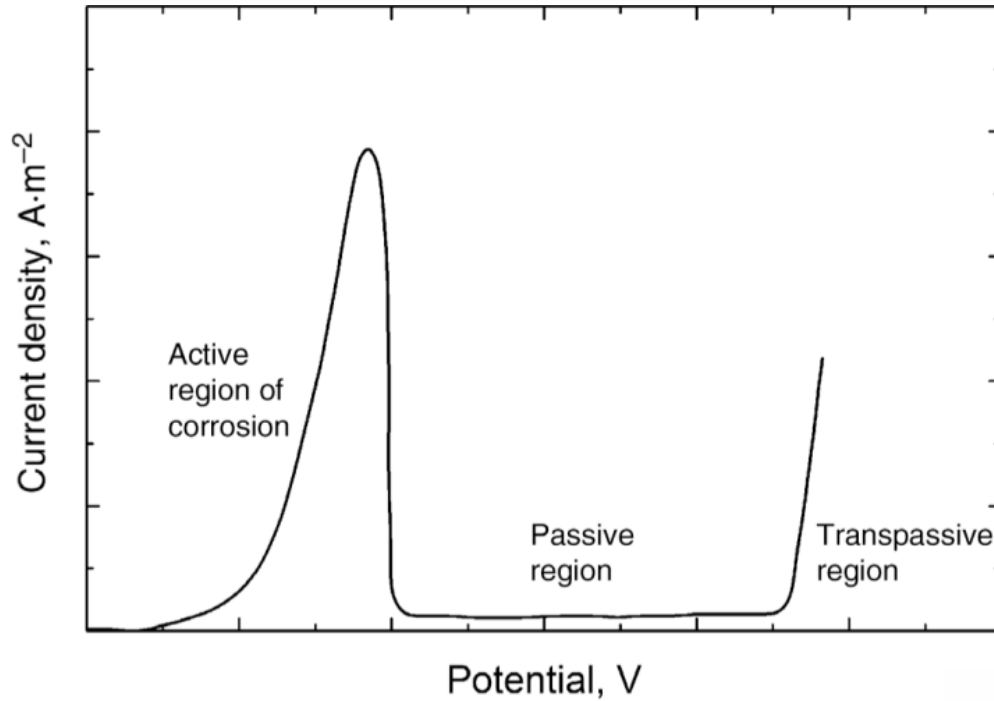
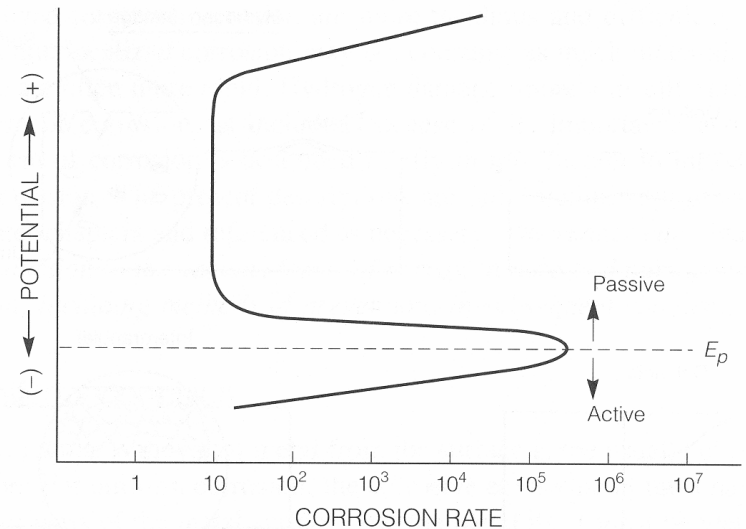


Figure 16.3 Current-voltage curve illustrating passivation.

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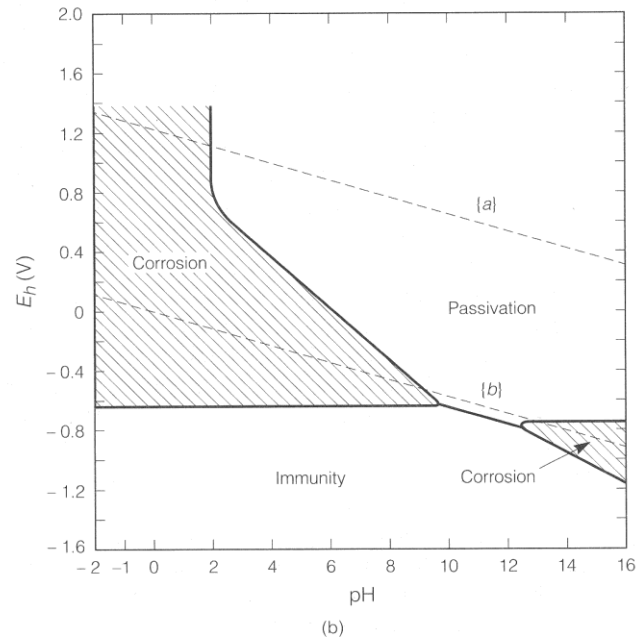
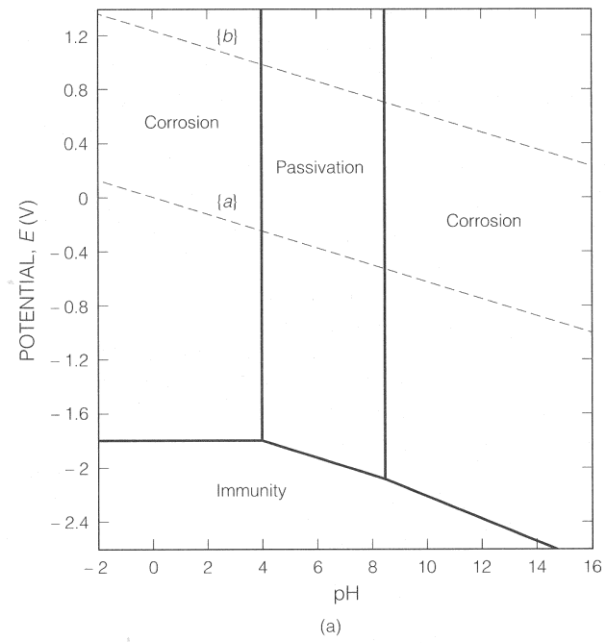


FIGURE 2.5 Pourbaix diagrams for (a) aluminum and (b) iron. All dissolved species at activities of 10^{-6} g-equiv/L.

Corrosion rate for uniform corrosion

- For a uniformly corroding surface, the potential naturally moves to the **corrosion potential**, where the anodic and cathodic currents are equal

Evans diagrams

- Tafel plot can be used to illustrate and to help us analyze uniform corrosion
→ I-V curves for both the anodic and cathodic reactions on the same plot

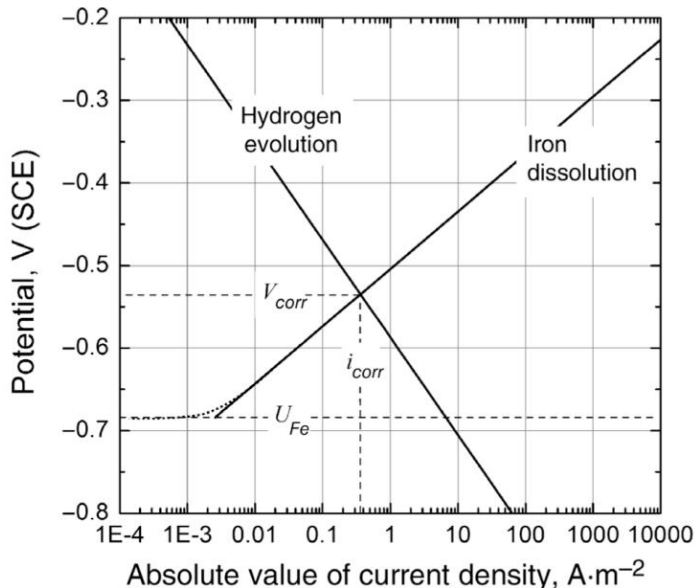


Figure 16.4 Evans diagram showing iron dissolution and hydrogen evolution (pH=0.9, 1 N HCl, deaerated). *Source:* Adapted from R. J. Chin and K. Nobe 1972, J. Electrochem. Soc., 119, 1457.

- The cathodic evolution of hydrogen and anodic dissolution of iron using Tafel expressions

- **V_{corr} (or E_{corr}): corrosion potential**
→ zero net current (since the rates of the two reactions are balanced),
i_{corr} (corrosion current)
→ the rate at which the metal will dissolve

Using Butler-Volmer equation and Tafel kinetics

$$\begin{aligned}
 V_{corr} &= \frac{\alpha_c U_c + \alpha_a U_a}{\alpha_a + \alpha_c} + \frac{RT}{F(\alpha_a + \alpha_c)} \ln \left(\frac{i_{o,c}}{i_{o,a}} \right) \\
 &= \frac{b_a U_c + |b_c| U_a}{b_a + |b_c|} + \frac{b_a |b_c|}{\ln(10)(b_a + |b_c|)} \log \left(\frac{i_{o,c}}{i_{o,a}} \right),
 \end{aligned}
 \tag{16.5}$$

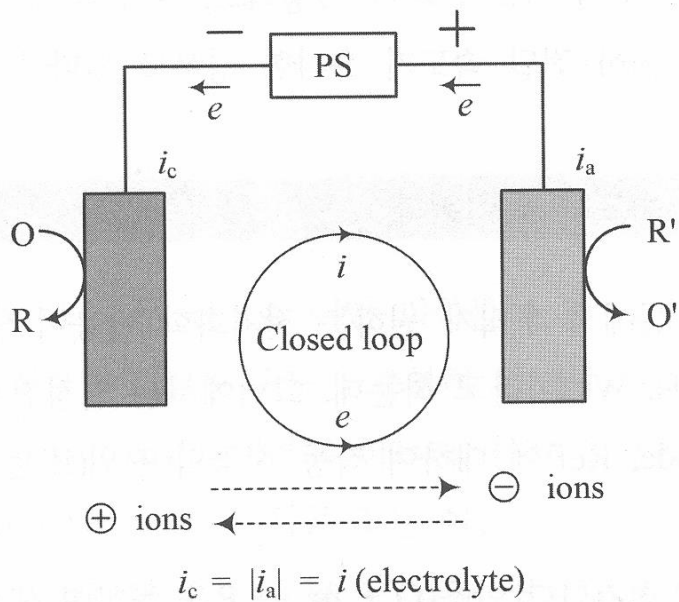
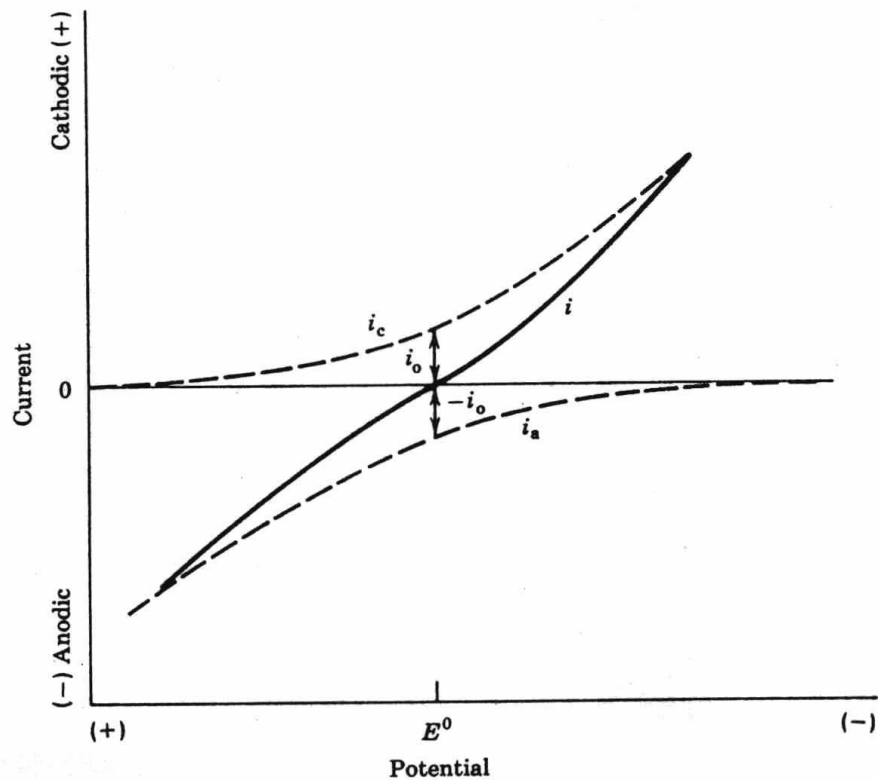
$$i_{corr} = i_{o,a} \left(\frac{\alpha_c}{\alpha_a + \alpha_c} \right) i_{o,c} \left(\frac{\alpha_a}{\alpha_a + \alpha_c} \right) \exp \left[\frac{F}{RT} \frac{\alpha_c \alpha_a}{\alpha_a + \alpha_c} (U_c - U_a) \right],
 \tag{16.6a}$$

$$= i_{o,a} \left(\frac{b_a}{b_a + |b_c|} \right) i_{o,c} \left(\frac{|b_c|}{b_a + |b_c|} \right) \exp \left[\frac{\ln(10)}{b_a + |b_c|} (U_c - U_a) \right],
 \tag{16.6b}$$

where b is the Tafel slope introduced in Chapter 3, which is negative for the cathodic reaction:

$$|b| = \frac{(\ln 10)RT}{\alpha F} = \frac{2.303RT}{\alpha F}.
 \tag{3.25}$$

From ch.13 Electrodeposition



I-V: Electroless deposition

■ Electrochemical Point of View: Mixed Potential Theory (혼성전위)

- Mixed potential theory: overlap of two independent electrochemical reaction

Metal ion reduction



$$i_c = f(E)$$

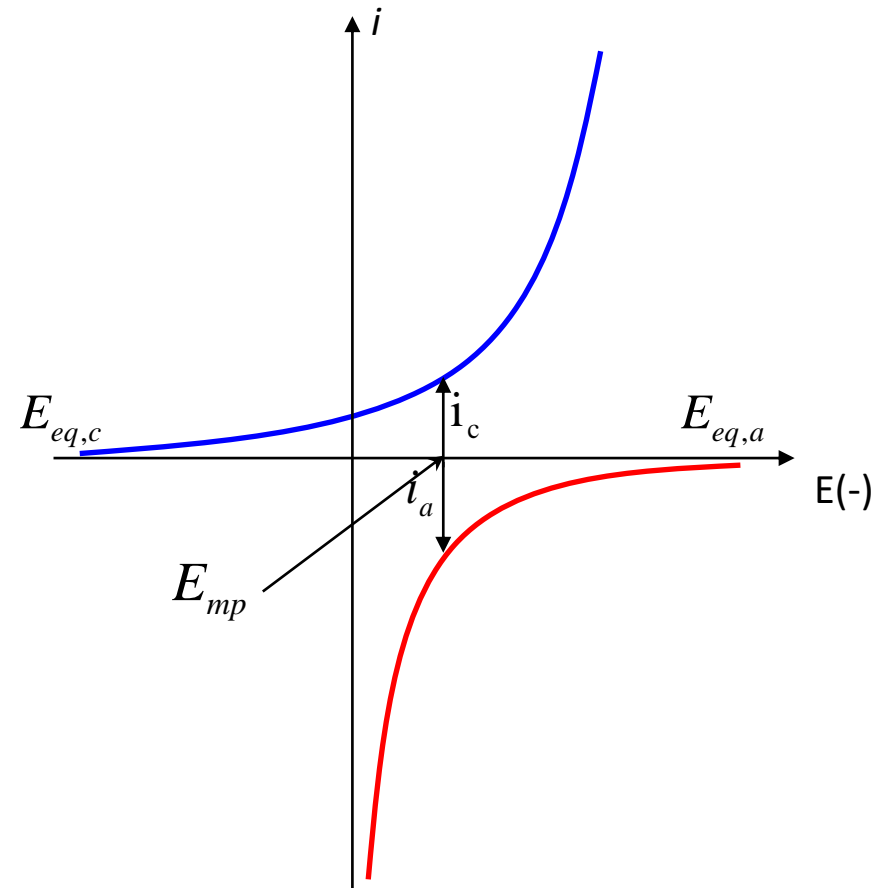
Reducing agent oxidation



$$i_a = f(E)$$

- The potential where the total current becomes 0 is the mixed potential.

$$i_{total} = i_c + i_a$$



Electroless deposition

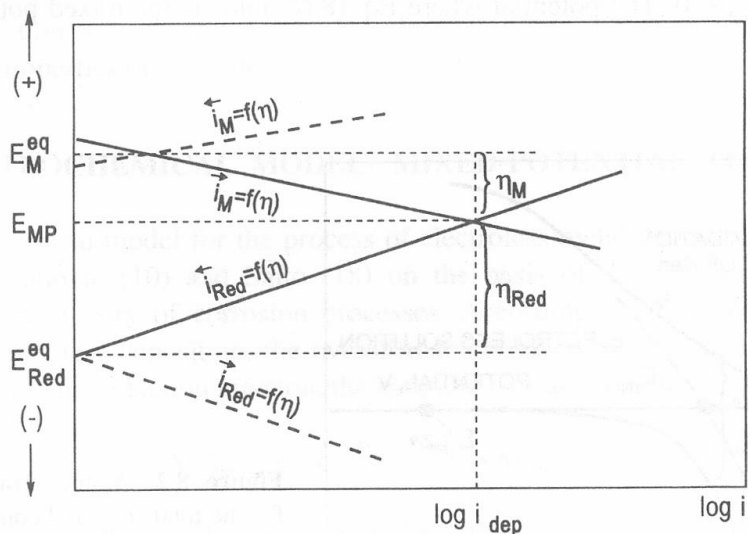


Figure 8.3. Evans diagram of current–potential curves for the system with two different simultaneous electrochemical reactions. Kinetic scheme: Eqs.(8.4) and (8.5).

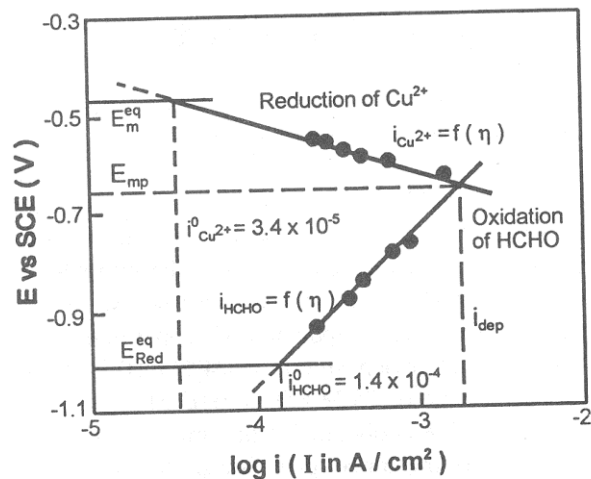
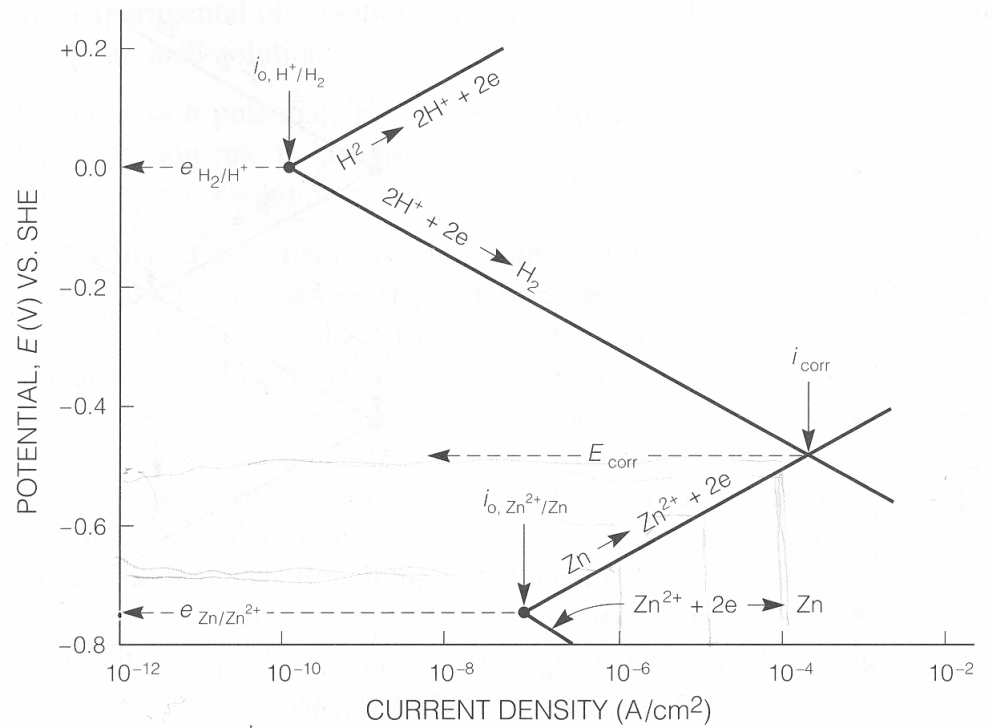
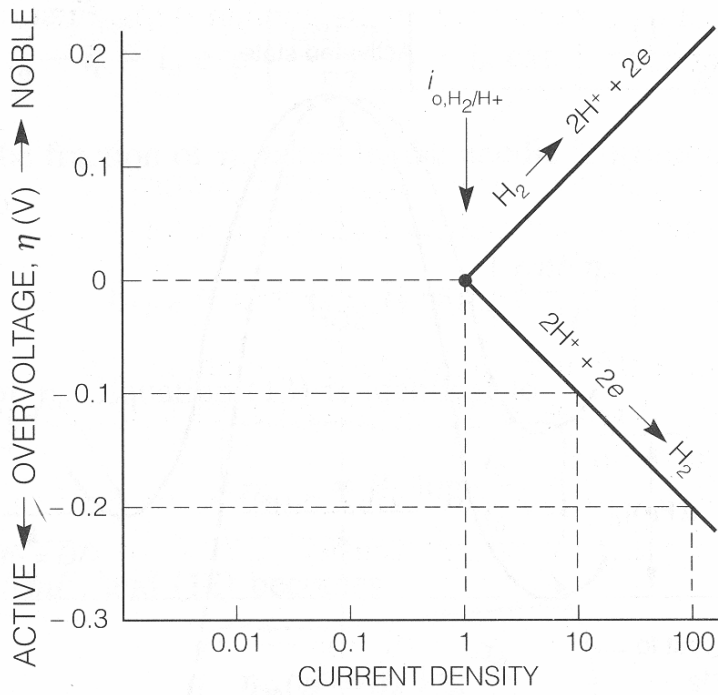


Figure 8.4. Current–potential curves for reduction of Cu^{2+} ions and for oxidation of reducing agent Red, formaldehyde, combined into one graph (Evans diagram). Solution for the Tafel line for the reduction of Cu^{2+} ions—0.1 M CuSO_4 , 0.175 M EDTA, pH 12.50, $E_{\text{eq}}(\text{Cu}/\text{Cu}^{2+}) = -0.47$ V versus SCE; for the oxidation of formaldehyde—0.05 M HCHO and 0.075 M EDTA, pH 12.50, $E_{\text{eq}}(\text{HCHO}) = -1.0$ V versus SCE; temperature 25°C ($\pm 0.5^\circ\text{C}$). (From Ref. 10, with permission from the American Electroplaters and Surface Finishers Society.)



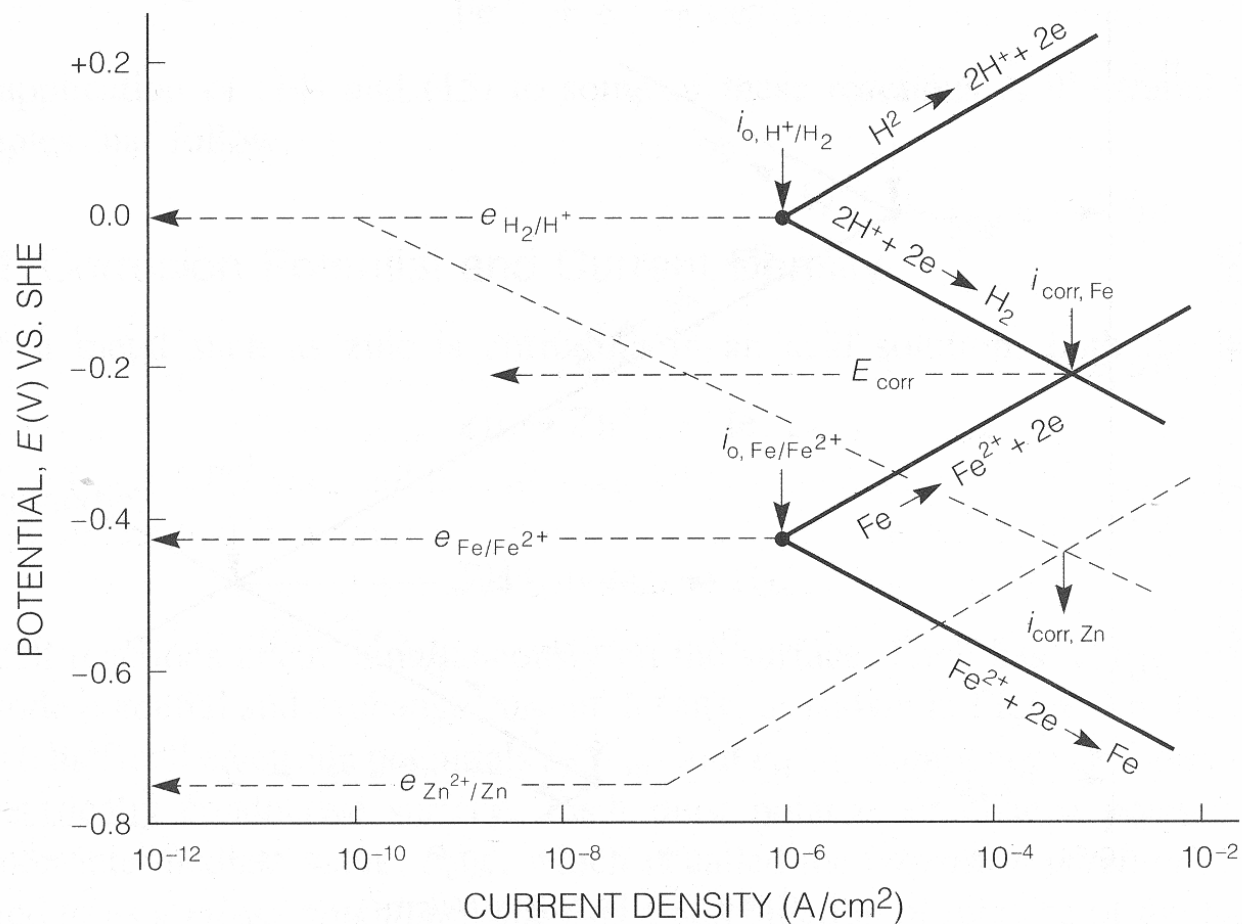


FIGURE 3.8 Comparison of electrochemical parameters for iron and zinc in acid solution, demonstrating the importance of i_o on determination of corrosion rates. Dashed lines represent lines from Figure 3.7 superimposed for comparison.

Experimental measurement of corrosion

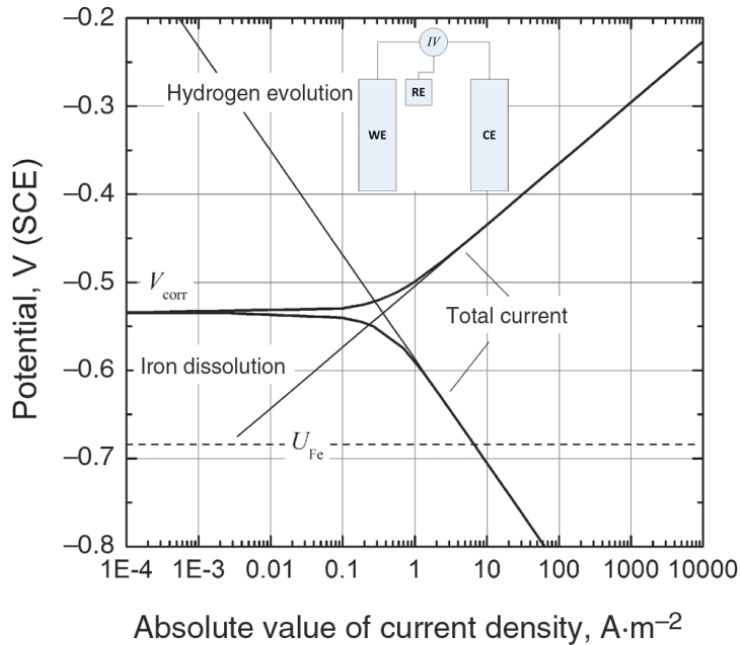


Figure 16.5 Relationship between the total current and the anodic and cathodic currents for a corrosion sample at different potentials.

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$$i = i_{corr} \left[\exp \frac{\alpha_a F}{RT} (V - V_{corr}) - \exp \frac{-\alpha_c F}{RT} (V - V_{corr}) \right]. \quad (16.7a)$$

$$= i_{corr} \left[\exp \frac{\ln(10)}{b_a} (V - V_{corr}) - \exp \frac{\ln(10)}{b_c} (V - V_{corr}) \right]. \quad (16.7b)$$

The limitations of Equation 16.7 should be considered when fitting data to this expression. Also, remember that this type of analysis is strictly valid only for systems undergoing uniform corrosion of the surface.

Equation 16.7 is the basis for a commonly used corrosion technique called *linear polarization resistance*. Linearization of Equation 16.7b about the corrosion potential yields

$$i = \frac{i_{corr} \ln(10)(b_a + |b_c|)}{b_a |b_c|} (V - V_{corr}) = \frac{1}{R_p} (V - V_{corr}). \quad (16.8)$$

R_p is defined as the polarization resistance:

$$\frac{b_a |b_c|}{i_{corr} \ln(10)(b_a + |b_c|)} \equiv R_p.$$

The units for the polarization resistance are $\Omega \cdot \text{m}^2$; alternatively, the use of the current instead of the current density would yield a resistance in ohms. If the Tafel coefficients are known, then the corrosion current density (or current) can be calculated from the polarization resistance as follows:

$$i_{corr} = \frac{b_a |b_c|}{\ln(10)(b_a + |b_c|) R_p}. \quad (16.9)$$

Mass transfer effects

- Mass transfer effect: e.g. oxygen reduction → limiting current

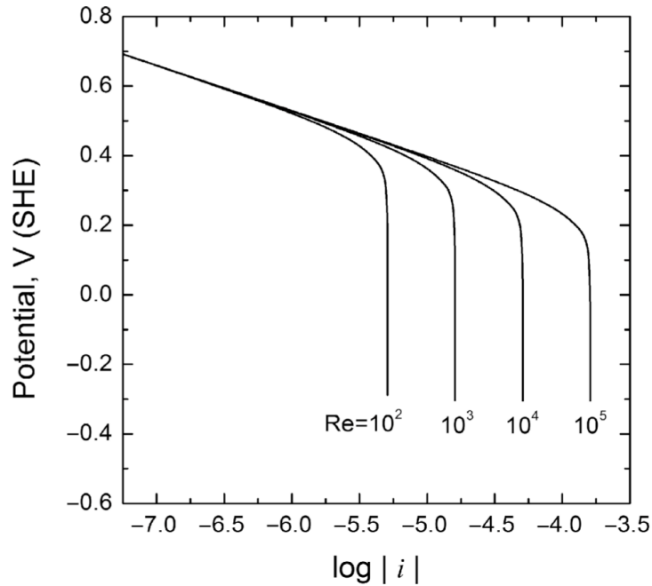


Figure 16.6 Oxygen reduction current at different flow rates for fully developed laminar flow over a flat plate.

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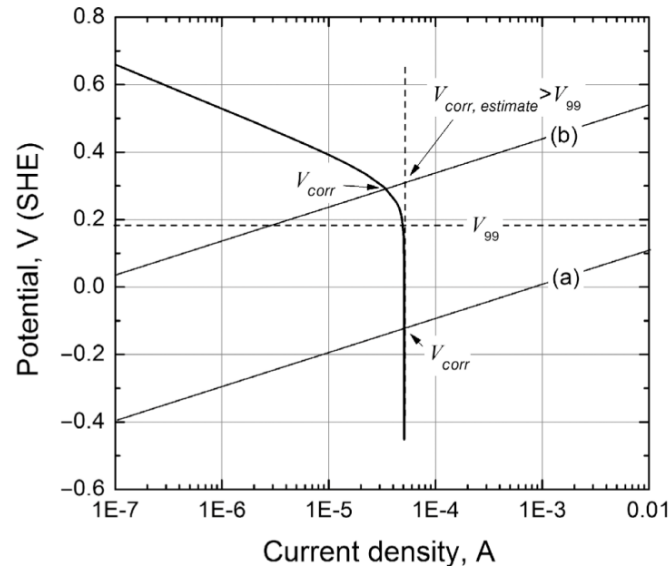


Figure 16.7 Illustration of corrosion potential determination for a cathodic reaction that is mass-transfer limited.

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Two anodic curves

(a) One that intersects in the mass transfer limited region

(b) One that intersects in the region where rate effects also contribute

Multiple reactions

- Multiple cathodic reactions (multiple anodic reactions later)

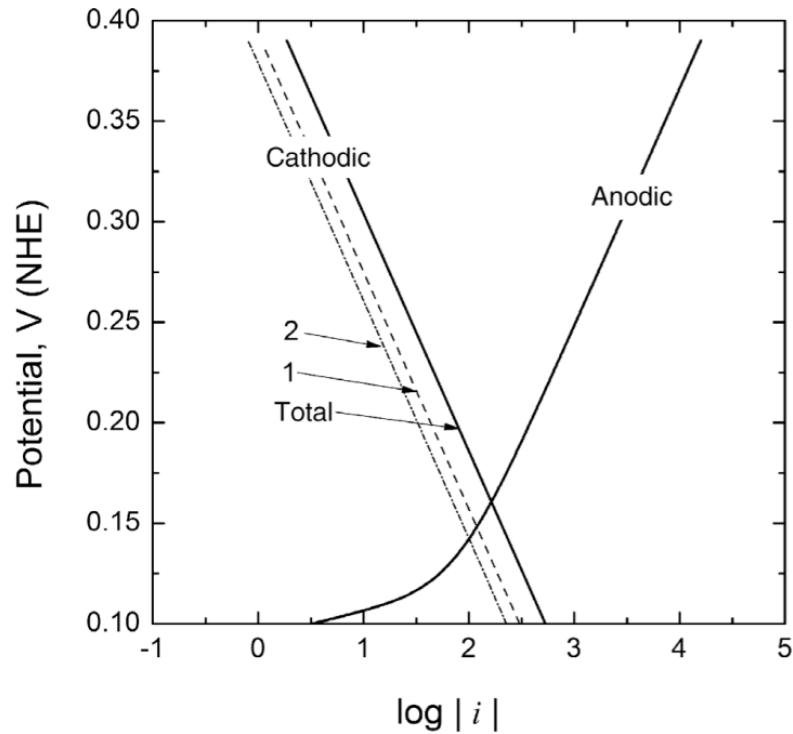


Figure 16.8 Multiple cathodic reactions for the corrosion of copper metal by copper(II) ions in acid.

Localized corrosion

- Localized corrosion is very different from the uniform corrosion (Fig.16.1)

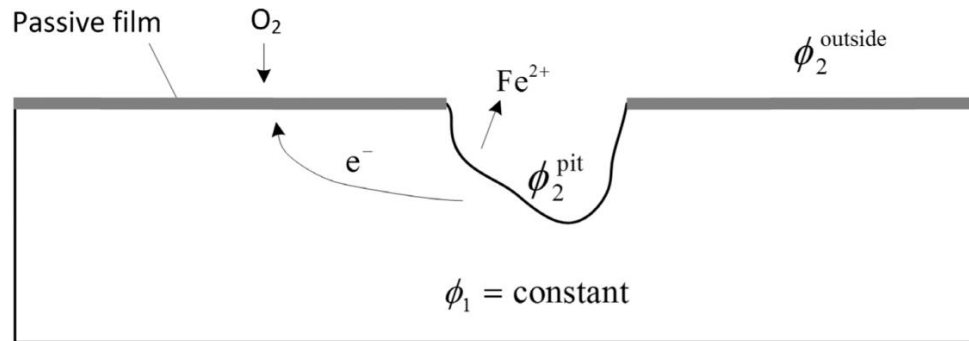


Figure 16.9 Schematic diagram of corrosion pit. Current flows from the inside of the pit to the outside, driven by a potential difference in solution. The potential is greater in the pit than outside the pit.

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Galvanic corrosion

- Galvanic corrosion is a form of localized corrosion that can occur when two different metals are in contact and exposed to a common electrolyte
- e.g. Zn + Fe

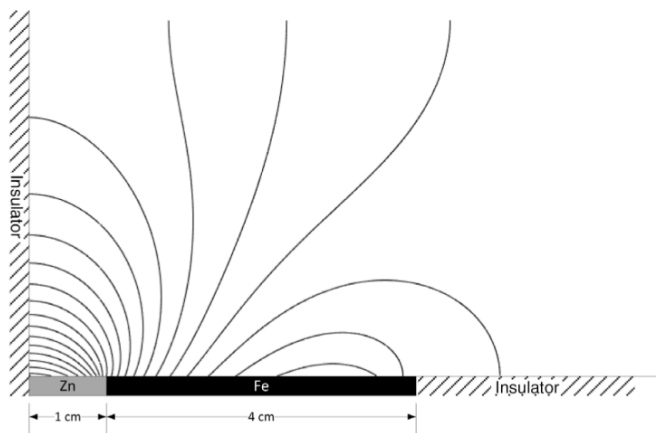


Figure 16.10 Potential field for galvanic coupling of Zn and Fe.
 Source: Data adapted from Abootalebi 2010.

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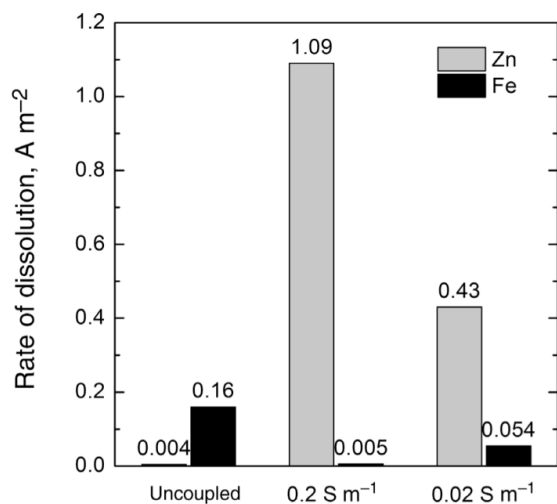


Figure 16.11 Average dissolution rates: (1) without galvanic coupling, (2) coupled with a solution conductivity of 0.2 S m^{-1} , and (3) coupled with a solution conductivity of 0.02 S m^{-1} .

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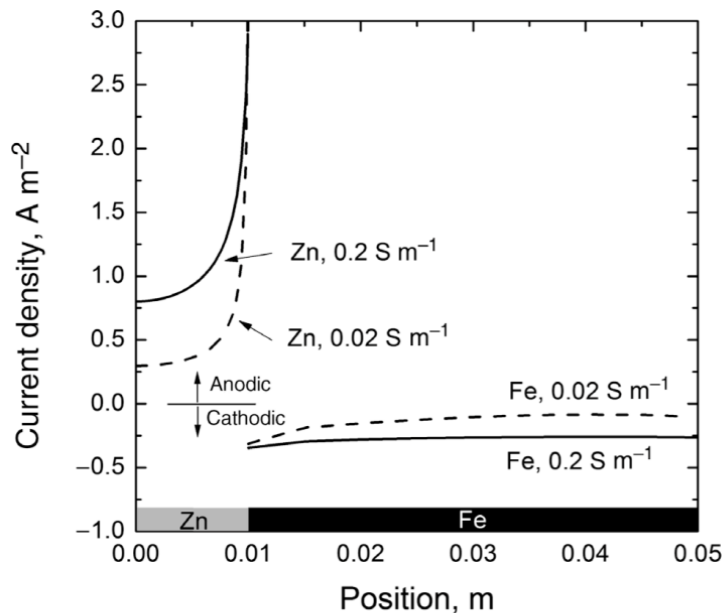


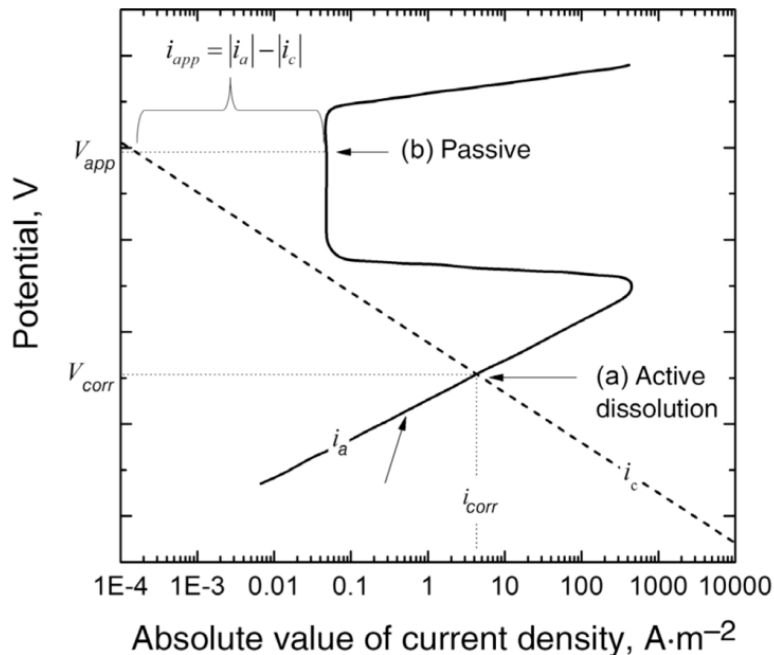
Figure 16.12 Local current density for galvanically coupled Zn and Fe at solution conductivities of 0.2 and 0.02 S m^{-1} .

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Corrosion protection

- Cathodic protection, anodic protection, electrophoretic coating (or e-coating)

Anodic protection: much less commonly than cathodic protection
-(i) Increasing potential in the anodic direction to passive region

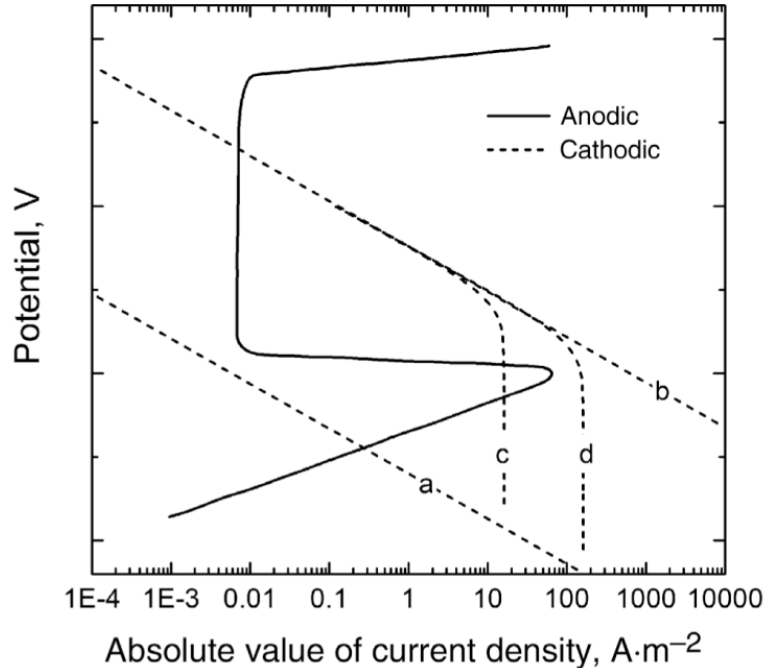


- anodic corrosion current is reduced by two orders
- Cathodic current is reduced more

Figure 16.13 Illustration of using anodic protection to reduce the rate of corrosion for metals that passivate.

Anodic protection

-(ii) Changing the cathodic reaction so that the cathodic curve intersects the anodic curve in the passive region



- Changing oxygen concentration

Figure 16.14 Importance of mass transfer on passivating the metal.

Cathodic protection

- used to protect sea vessels, oil platforms, ocean structures, buried pipelines, storage tanks and so on
- lower the potential of the metal in the cathodic direction to decrease its dissolution rate → to reduce or eliminate corrosion → ideally to the immunity area (Pourbaix diagram)

- Two ways for cathodic protection
 - Use of a sacrificial anode
 - Use of an external power supply to apply a cathodic current to the metal to be protected

V_1 : lower dissolution

V_2 : eliminate dissolution

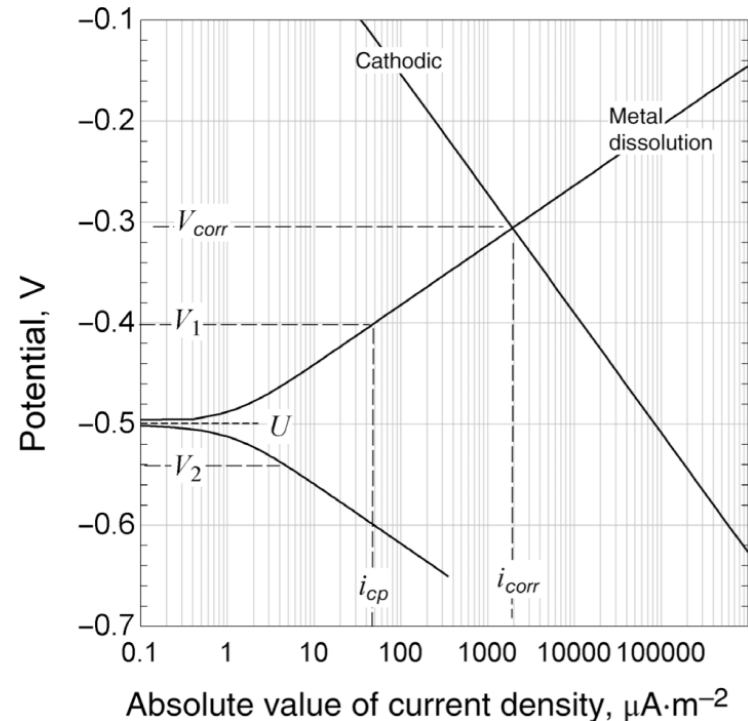


Figure 16.15 Evans diagram showing corrosion potential and the reduction of corrosion current with a reduction in potential.

(i) Use of a sacrificial anode

- coupling the metal to be protected with a more active metal, that is, one with a lower equilibrium potential

→ sacrificial metal (by galvanic corrosion)



Figure 16.16 Use of sacrificial anodes to protect against ship corrosion.

(i) Use of impressed current cathodic protection (ICCP)

- the external electrode is the anode (corroded) and the structure to be protected is the cathode

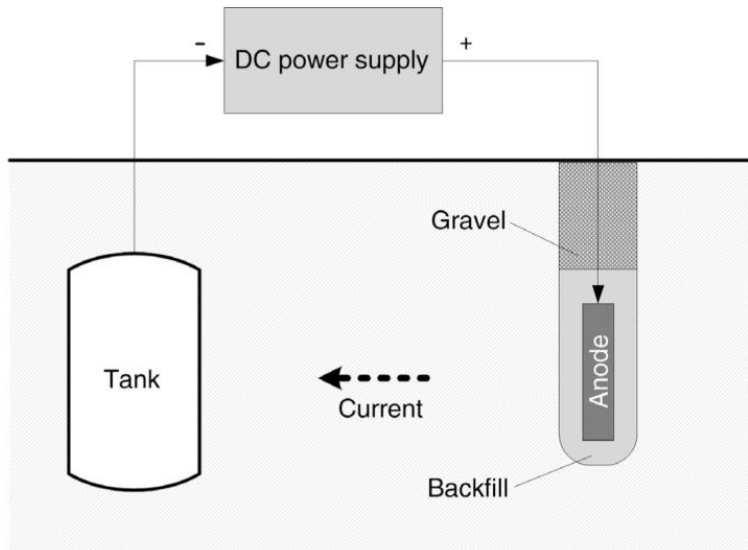


Figure 16.17 Impressed-current cathodic protection used to reduce corrosion of an underground tank.

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Table 16.1 Characteristics of ICCP Anodes for Seawater Applications

Material	Current density range ($\text{A}\cdot\text{m}^{-2}$)	Consumption rate $\text{kg}\cdot\text{A}^{-1}\cdot\text{yr}^{-1}$
Scrap steel	<1	9
Silicon iron	10–30	0.2–0.5
Graphite	10–20	0.3–0.5
Lead silver	200–300	0.1
Platinized anodes	250–2000	10^{-5}
Mixed metal oxide	250–600	$0.5\text{--}4 \times 10^{-6}$

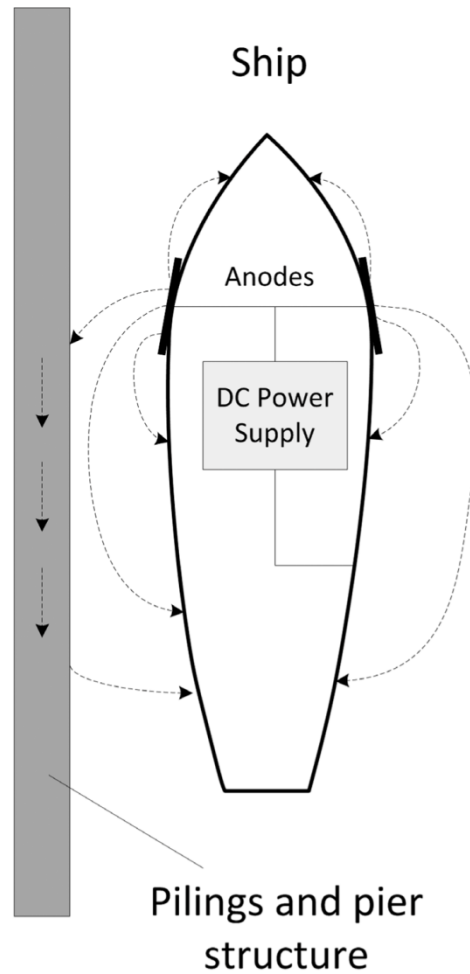


Figure 16.18 Stray currents lead to increased corrosion of pier.

TABLE 1.1 Galvanic Series in Seawater

<i>Cathodic (noble)</i>
↑
platinum
gold
graphite
titanium
silver
zirconium
AISI Type 316, 317 stainless steels (passive)
AISI Type 304 stainless steel (passive)
AISI Type 430 stainless steel (passive)
nickel (passive)
copper-nickel (70-30)
bronzes
copper
brasses
nickel (active)
naval brass
tin
lead
AISI Type 316, 317 stainless steels (active)
AISI Type 304 stainless steel (active)
cast iron
steel or iron
aluminum alloy 2024
cadmium
aluminum alloy 1100
zinc
magnesium and magnesium alloys
↓
<i>Anodic (active)</i>

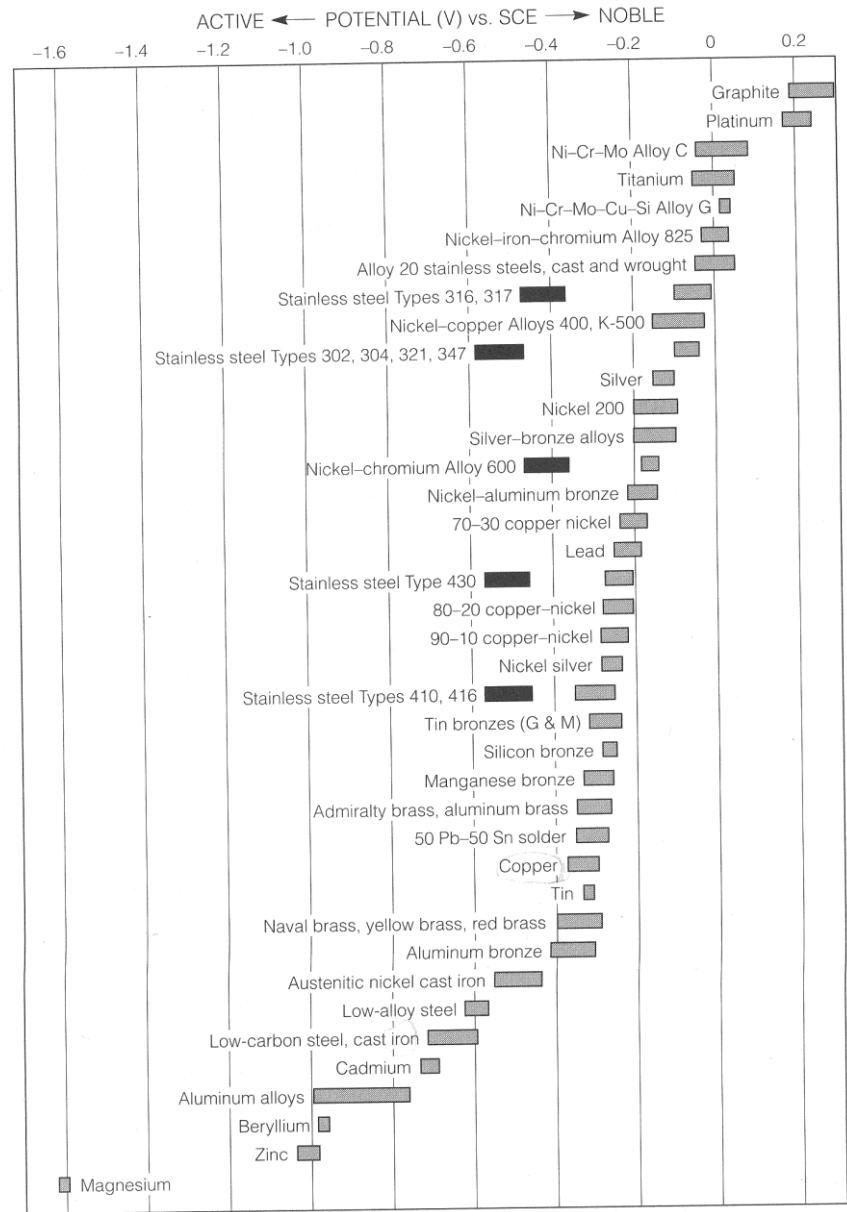
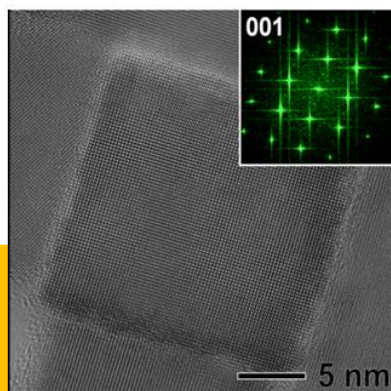


FIGURE 6.1 Galvanic Series for seawater. Dark boxes indicate active behavior for active-passive alloys. (From H. P. Hack, *Metals Handbook, Vol. 13, Corrosion, 9th ed.*, ASM, Metals Park, OH, p. 234, 1987. Reprinted by permission, ASM International.)

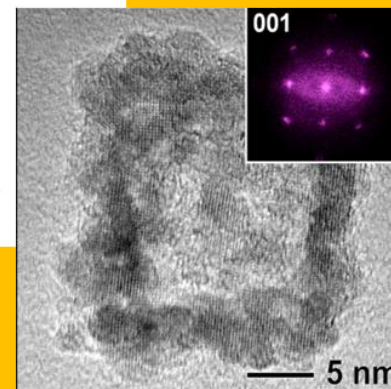
Galvanic Replacement Reactions in Metal Oxide Nanocrystals

Galvanic corrosion

Manganese Oxide



Iron Oxide

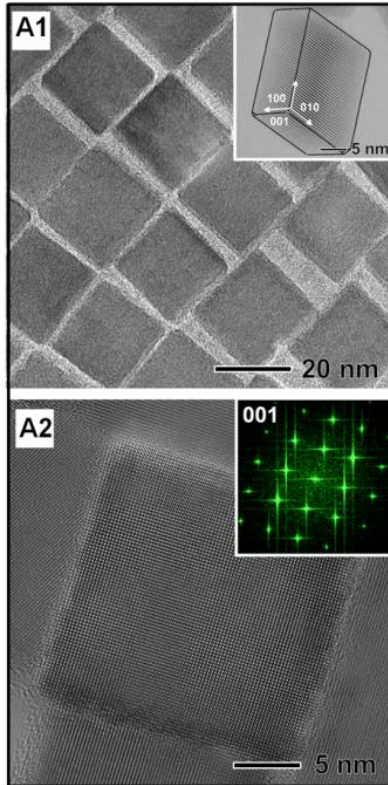


The galvanic replacement reaction is the most versatile method of preparing hollow metallic nanostructures with controllable pore structures and compositions (1–7). These

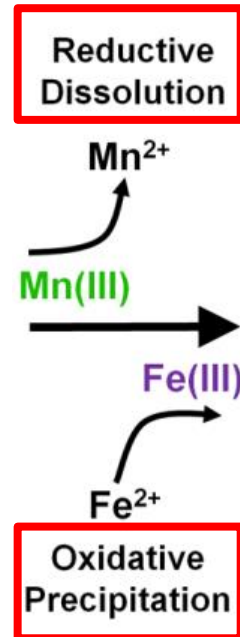
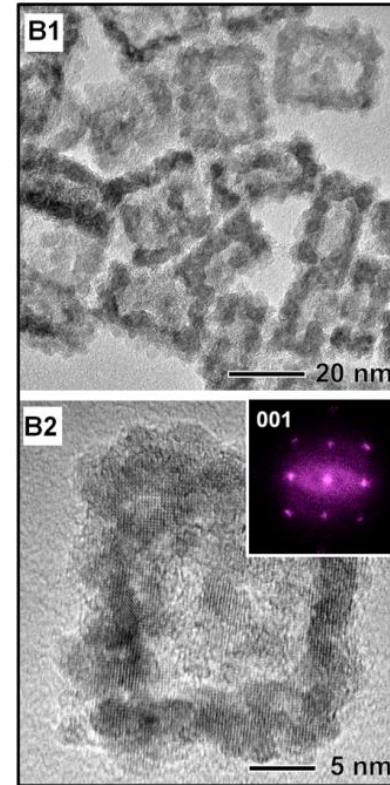
reactions involve a corrosion process that is driven by the difference in the electrochemical potentials of two metallic species. The hollow interior is generated from the oxidative dissolution of

Nanoparticle synthesis using electrodeposition/corrosion

Mn₃O₄ Nanocrystals



γ-Fe₂O₃ Nanocages



Standard reduction potential(V)	Multivalent ion pairs
1.87	Co ²⁺ /Co ³⁺
1.67	Pb ²⁺ /Pb ⁴⁺
1.61	Ce ³⁺ /Ce ⁴⁺
1.51	Mn ²⁺ /Mn ³⁺
0.77	Fe ²⁺ /Fe ³⁺
0.16	Cu ⁺ /Cu ²⁺
0.15	Sn ²⁺ /Sn ⁴⁺
0.00	Ti ³⁺ /Ti ⁴⁺
-0.40	Cr ²⁺ /Cr ³⁺

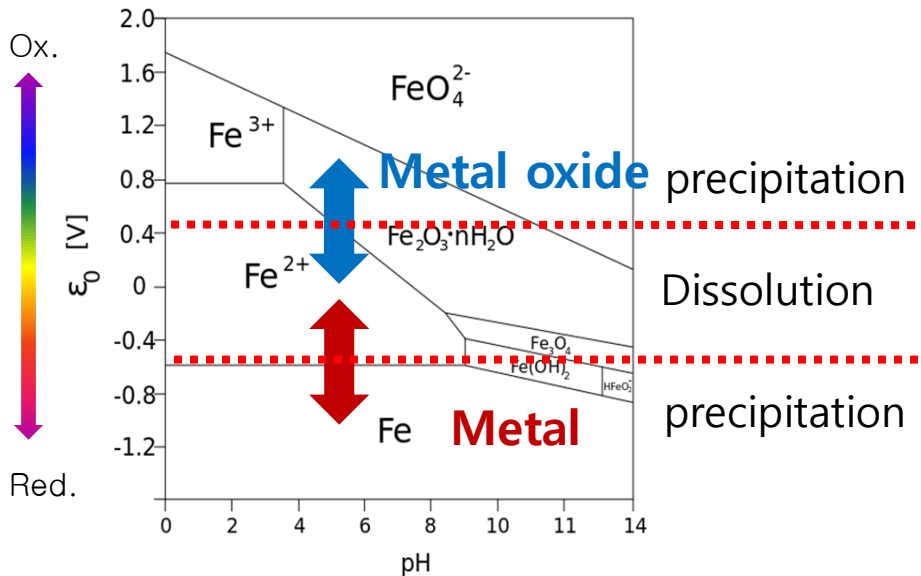
- Mn₃O₄ nanocrystals were reacted with Fe(II) perchlorate in a suspension.
- The lattice orientation of original Mn₃O₄ was preserved during the complete transformation.

Galvanic Replacement Reactions in Metal Oxide Nanocrystals

Metal Nanocrystal vs. Metal Oxide Nanocrystal

	Metal nanocrystal	Metal oxide nanocrystal
Representative equation	$3\text{Ag} + \text{HAuCl}_4 \rightarrow 3\text{AgCl} + \text{Au} + \text{HCl}$	$\text{Mn}_3\text{O}_4 + 2\text{Fe}^{2+} + 2\text{H}^+ \rightarrow \gamma\text{-Fe}_2\text{O}_3 + 3\text{Mn}^{2+} + \text{H}_2\text{O}$
Redox pair	Metal vs Metal ion	Metal ion vs Metal ion
Reaction mechanism	Oxidative dissolution and reductive precipitation	Reductive dissolution and oxidative precipitation

Pour-baix diagram

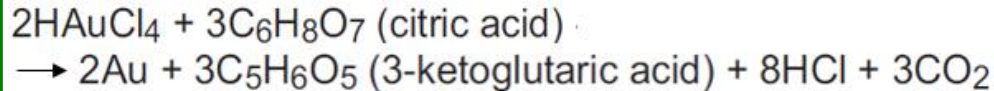


standard reduction potentials of multivalent ions

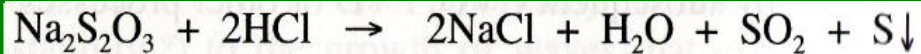
Standard reduction potential (V)	Multivalent ion pairs
1.87	$\text{Co}^{2+}/\text{Co}^{3+}$
1.67	$\text{Pb}^{2+}/\text{Pb}^{4+}$
1.61	$\text{Ce}^{3+}/\text{Ce}^{4+}$
1.51	$\text{Mn}^{2+}/\text{Mn}^{3+}$
0.77	$\text{Fe}^{2+}/\text{Fe}^{3+}$
0.16	$\text{Cu}^+/\text{Cu}^{2+}$
0.15	$\text{Sn}^{2+}/\text{Sn}^{4+}$
0.00	$\text{Ti}^{3+}/\text{Ti}^{4+}$
-0.40	$\text{Cr}^{2+}/\text{Cr}^{3+}$

Nanoparticle synthesis

2. *Gold sol.* Add 1 cm³ of 1% HAuCl₄.3H₂O to 100 cm³ of distilled water. Bring to the boil and add 2.5 cm³ of 1% sodium citrate. Keep the solution just boiling. A ruby red gold sol forms after a few minutes.



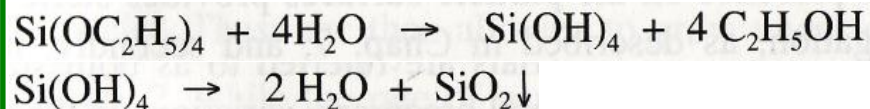
3. *Sulphur sol.* Mix equal volumes of aqueous solutions (10⁻³ to 5 × 10⁻³ mol dm⁻³) of Na₂S₂O₃ and HCl.



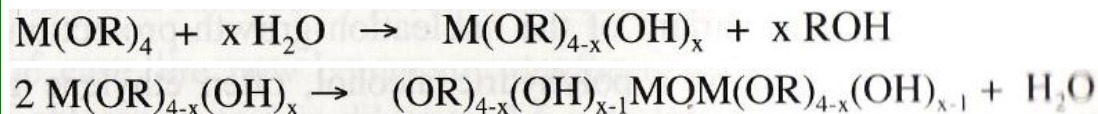
4. *Hydrous iron(III) oxide sol.* Add, with stirring, 2 cm³ of 30% FeCl₃(aq) to 500 cm³ of boiling distilled water. A clear reddish-brown dispersion is formed.



5. Hydrolysis + condensation of inorganic alkoxide precursors:



TEOS (tetraethyloxysilane) or
TMOS(tetramethyloxysilane)



hydrolysis,
condensation

M = metal



SCHEDULES (could be modified)

- Introduction and basic principles (ch.1) (1 week)
- Cell potential and thermodynamics (ch.2) (2 week)
- Electrochemical kinetics (ch.3) (3 week)
- Transport (ch.4) (4 week)
- Electrode structures and configurations (ch.5) (5 week)
- Analysis of electrochemical systems (ch.6) (6 week)
- Batteries 1, 2 (ch.7,8) (7,8 weeks)
- Fuel cells 1, 2 (ch.9,10) (9,10 weeks)
- Electrochemical capacitors (ch.11) (11 week)
- Energy storage & conversion (ch.12) (11 week)
- Electrodeposition (ch.13) (12 week)
- Industrial electrolysis (ch.14) (13 week)
- Photoelectrochemical cells (ch.15) (14 week)
- Corrosion (ch.16) (15 week)

Final Exam (June 13)

Figures in Lecture Notes!!

TEXTBOOK

Thomas F. Fuller, John N. Harb, *Electrochemical Engineering*, Wiley, 2018.
(It is recommended to solve the textbook example(illustration))

REFERENCES (참고문헌들)

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Allen J. Bard, Larry R. Faulkner, *Electrochemical Methods*, Wiley, 2001 (Korean Ver.)

GRADING (B⁺ & above ~ 80%, B⁰ & below ~ 20%, or Departmental guide)

Midterm Exam 40%, Final Exam 40%, Homeworks & Attendance 20 %

LECTURE ROOM & TIME: Rm #302-409, 11:00-12:15 Mon. & Wed.

OFFICE HOUR: Rm #302-721, 13:00-16:00 Mon. & Wed.

•TA: Geumbi Na(나금비) Rm #302-1007, 02-880-9123, mstu2025@snu.ac.kr

Objectives of the Electrochemical Energy Engineering Lecture

전기화학에너지공학 강의의 목표

- ✓ Cultivation of basic knowledge about electrochemistry.
 - ✓ Cultivation of basic knowledge of electrochemical engineering or industrial application.
 - ✓ Exploring the applicability of research through the electrochemical reaction of oxidation and reduction.
 - ✓ Gain insight into R&D through the history of electrochemistry.
 - ✓ Securing synergy by exploring the possibility of joint research with other fields.
-
- ✓ **전기화학(Electrochemistry)에 대한 기초적 지식 함양**
 - ✓ **전기화학의 공학 혹은 산업적 응용 기초지식 함양**
 - ✓ **산화와 환원의 전기화학반응을 통한 연구에의 적용 가능성 탐색**
 - ✓ **전기화학 역사를 통한 연구개발의 통찰력 확보**
 - ✓ **타분야와의 공동연구 가능성 모색을 통한 시너지 확보**

