Electrochemical Engineering
Lecture 05 Applications

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Battery Introduction

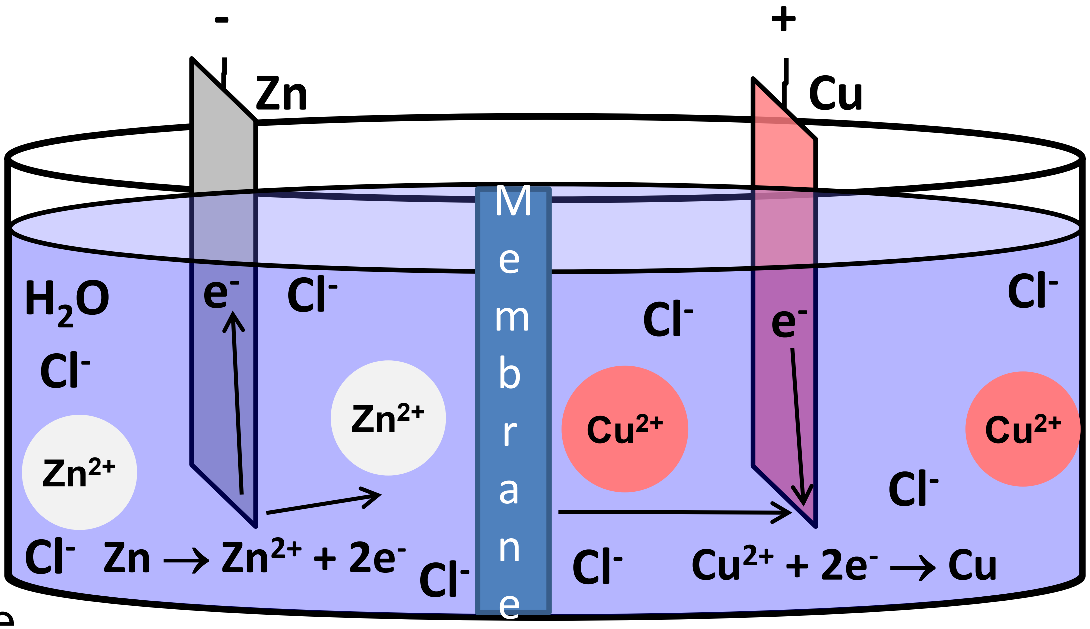
- **Electrochemical system that converts chemical energy in <u>solids</u> to <u>electricity</u> via electrochemical reactions (and vice versa for rechargeable battery)**
- **Example - Daniell cell (1839)**
  - Negative electrode

    $Zn^{2+} + 2e^- = Zn$

    $E^o(Zn^{2+}/Zn) = -0.763$ V vs. SHE
  - Positive electrode

    $Cu^{2+} + 2e^- = Cu$

    $E^o(Cu^{2+}/Cu) = 0.337$ V vs. SHE
  - Electrolyte, e.g., $ZnCl_2$ & $CuCl_2$ dissolved in water
  - Membrane separator/salt bridge



  - Overall/full cell reaction: $Zn + Cu^{2+} \underset{\text{Charge}}{\overset{\text{Discharge}}{\rightleftharpoons}} Zn^{2+} + Cu$
  - <u>(Standard)</u> **cell** potential $E_{cell}^o = E_{cat}^o - E_{an}^o = 0.337\text{ V} - (-0.763\text{ V}) = 1.10\text{ V}$
  - Nernst equation to correct for non-standard conditions

Battery Application & Rechargeability

❑Applications

- Power sources for mobile electronics, automobile, etc.
- Backup power source for critical systems: computer server, emergency lighting
- (Renewable) energy storage
- …

https://en.wikipedia.org/wiki/Electric_battery#/media/File:Batteries.jpg
https://www.powerbankexpert.com/how-long-do-cell-phone-batteries-last/
https://www.caranddriver.com/shopping-advice/a26310824/best-car-battery/

https://wonderfulengineering.com/what-is-an-ups/
https://www.abc.net.au/news/2020-09-02/tesla-battery-expanded-as-sa-energy-minister-lauds-benefits/12622382

❑Classification by recharge capability

- **Primary** battery - no recharging: alkaline (Zn-MnO_2)
- **Secondary** or **rechargeable** battery: lead acid, Ni-Cd, NiMH, lithium ion, etc.

Battery Classification by Chemistry & Mechanism

Classification by chemistry - Infinite possibilities

Lead-acid

Alkaline

Silver-zinc

Zinc-air

Ni-Cd battery

Nickel metal hydride (NiMH)

Lithium ion (LIB)

https://www.caranddriver.com/shopping-advice/a26310824/best-car-battery/
https://en.wikipedia.org/wiki/Alkaline_battery
https://biz.maxell.com/en/primary_batteries/sr_coin.html
https://en.wikipedia.org/wiki/Nickel%E2%80%93cadmium_battery
https://www.pololu.com/product/1003
https://www.powerbankexpert.com/how-long-do-cell-phone-batteries-last/

- Others
  - Na-S and Li-S
  - Li-O_2 (Li air)

Classification by mechanism

- Conversion/reconstruction - most batteries (e.g., lead acid, alkaline, Ni-Cd, NiMH)
- Insertion - LIB

Requirements for Batteries

- Fast electrode reactions
- Equilibrium cell potential >1V w/ working potential >~0.5V
- Good stability with little self-discharging or unwanted reactions with environment under typical packaging
- High power & energy density
- Low cost
- …
- A single battery excels in certain aspect(s) but lags in others

Lead-Acid Battery

❑Reactions involved for discharge

Negative electrode (anode):
$PbSO_4 + 2H^+ + 2e^- = Pb + H_2SO_4$
$E^o(PbSO_4/Pb) = -0.359$ V vs. SHE

Positive electrode (cathode):
$PbO_2 + H_2SO_4 + 2H^+ + 2e^- = PbSO_4 + 2H_2O$
$E^o(PbO_2/PbSO_4) = 1.691$ V vs. SHE

Overall/full cell reaction:
$Pb + PbO_2 + 2H_2SO_4 = 2PbSO_4 + 2H_2O$
$E^o_{cell} = E^o_{cat} - E^o_{an} = 1.691\text{ V} - (-0.359\text{ V}) = 2.05\text{ V}$

https://www.caranddriver.com/shopping-advice/a26310824/best-car-battery/

❑Practical considerations

- Microstructure optimization: dense Pb & PbO_2 give low current density
- Should NOT fully discharge, i.e., convert all Pb/PbO_2 to $PbSO_4$ as it is insulating!

Alkaline or Zn-MnO_2 Battery

https://en.wikipedia.org/wiki/Alkaline_battery

❑Reactions involved for discharge

Negative electrode (anode):
$Zn^{2+} + 2e^- = Zn$

Positive electrode (cathodic):
$MnO_2 + H_2O + e^- = MnOOH + OH^-$

Electrolyte:
$Zn^{2+} + 2NH_4Cl + 2OH^- = Zn(NH_3)_2Cl_2 + H_2O$

Overall/full cell reaction:
$Zn + 2MnO_2 + 2NH_4Cl = MnOOH + Zn(NH_3)_2Cl_2$

❑Practical considerations

MnO_2 not very conductive and must be mixed with graphite powder and electrolyte to achieve decent rate capability

Lithium Ion Battery (LIB)

❑Reactions involved for discharge

https://www.powerbankexpert.com/how-long-do-cell-phone-batteries-last/

Negative electrode (anode):
$6C + xLi^+ + xe^- = Li_xC_6$

Positive electrode (cathodic):
$xLi^+ + Li_{1-x}CoO_2 + xe^- = LiCoO_2$

Overall/full cell reaction:
$Li_xC_6 + Li_{1-x}CoO_2 = 6C + LiCoO_2$
$E^o_{cell} \approx 4.0$ V

❑Practical considerations

- High energy density and power density comparing with other batteries
- High efficiency and no memory effect
- Flammable electrolyte

OCV, Specific Capacity, & Specific Energy

❑Open circuit voltage (OCV)

If no internal shorting, $E_{OCV} \approx E_{eq}$

❑Specific capacity

Charge stored for an electrode or an electrochemical cell, given by:

$$SC_{theory} = \frac{nF}{\sum_i x_i M_i}$$

Unit: C/g, mA · h/g, **A·h/g …**

n	Number of electron transferred in the half cell or full cell reaction
F	Faraday constant
x_i	Stoichiometry coefficient for species i in the half cell or full cell reaction
M_i	Molar mass for species i in the half cell or full cell reaction

- Actual specific capacity values much lower due to nonactive but necessary components such as electrolyte, current collectors, and packaging

❑Specific energy

Energy stored for an electrode or an electrochemical cell, given by

$$SE_{theory} = SC_{theory} \cdot OCV = \frac{nF}{\sum_i x_i M_i} \cdot OCV$$

Unit: J/g, mA · h /kg, **W · h/g** …

SOC or SOD

❑State of charge (SOC)

$$SOC = \frac{available\ capacity\ remaining}{total\ capacity} \times 100\%$$

Related to state of discharge (SOD) or depth of discharge (DOD) $SOD = 100\% - SOC$

❑Cell OCV changes w/ SOD (or SOC)

- OCV decreases w/ SOD
- Exact behavior depends on Chemistry
- Example - Lead acid battery

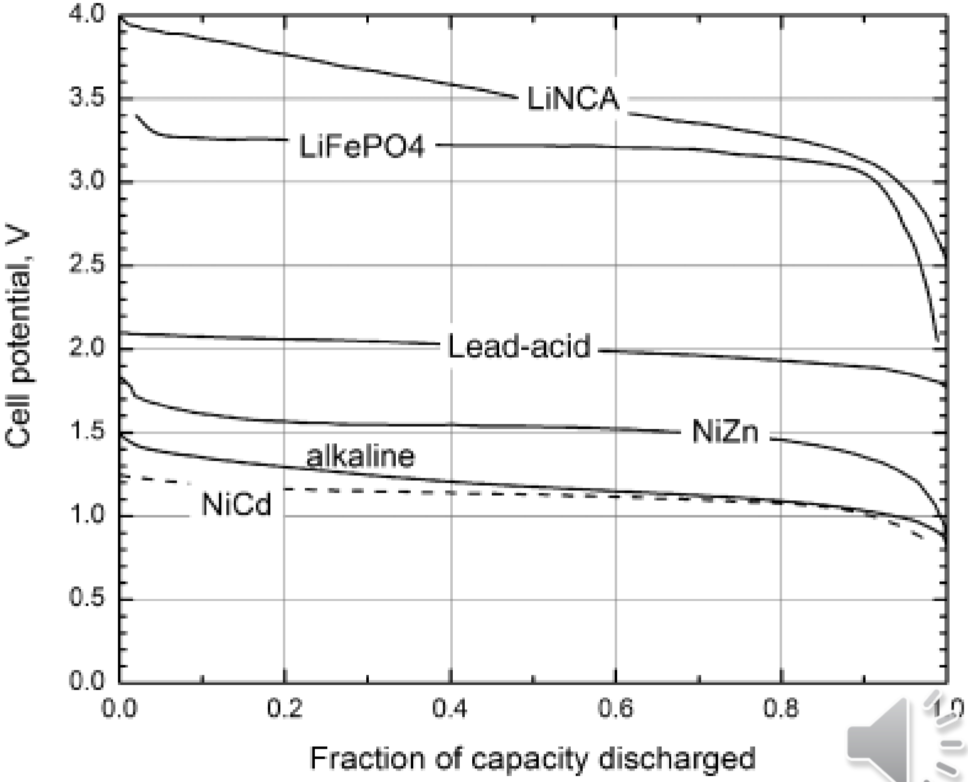
$Pb + PbO_2 + 2H_2SO_4 =$
$2PbSO_4 + 2H_2O$

$$E = E_{eq}^{o} - \frac{RT}{2F}\ln\left(\frac{a_{H_2O}}{a_{H_2SO_4}}\right)^2$$

- Battery discharge → SOD↑

$$a_{H_2SO_4}\downarrow;\ \ a_{H_2O} > a_{H_2SO_4}\ \&\ a_{H_2O}\uparrow$$

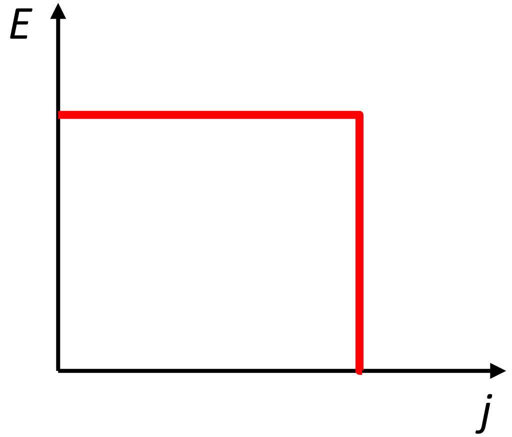
$$\rightarrow\ \ E\downarrow$$



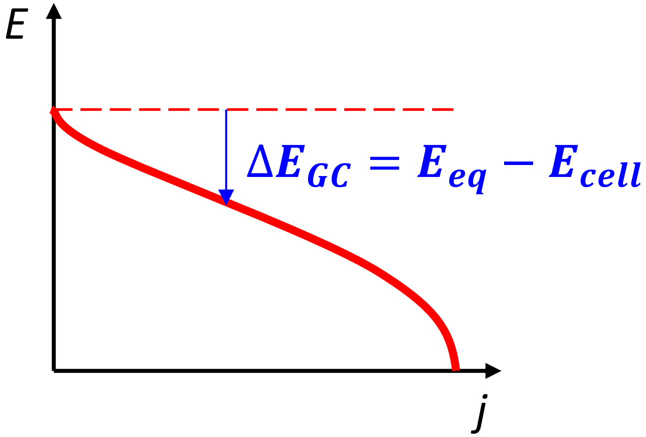
Fuller (2018), Fig. 7.5

Battery Discharge & Voltage Losses

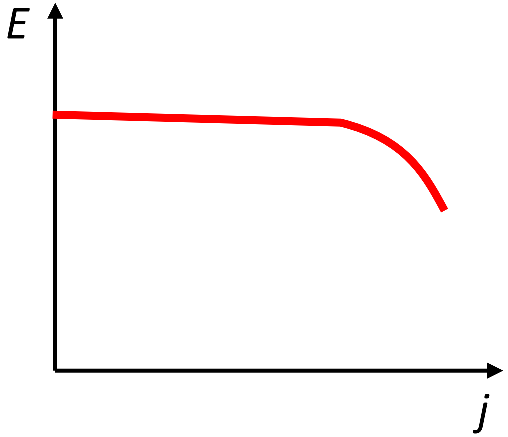
❑Different battery discharge characteristics



Ideal discharge with negligible internal resistance



Discharge with both activation & mass-transfer limitations



Discharge with fast electrode kinetics but mass-transfer limitation

❑Voltage losses

$$\Delta E_{GC} = E_{eq} - E_{cell} = IR_{int} = IR_{\Omega} + |\eta_{anode}| + |\eta_{cathode}| + |\eta_{concentration}|$$

Battery C Rate & Impact on Cell Potential

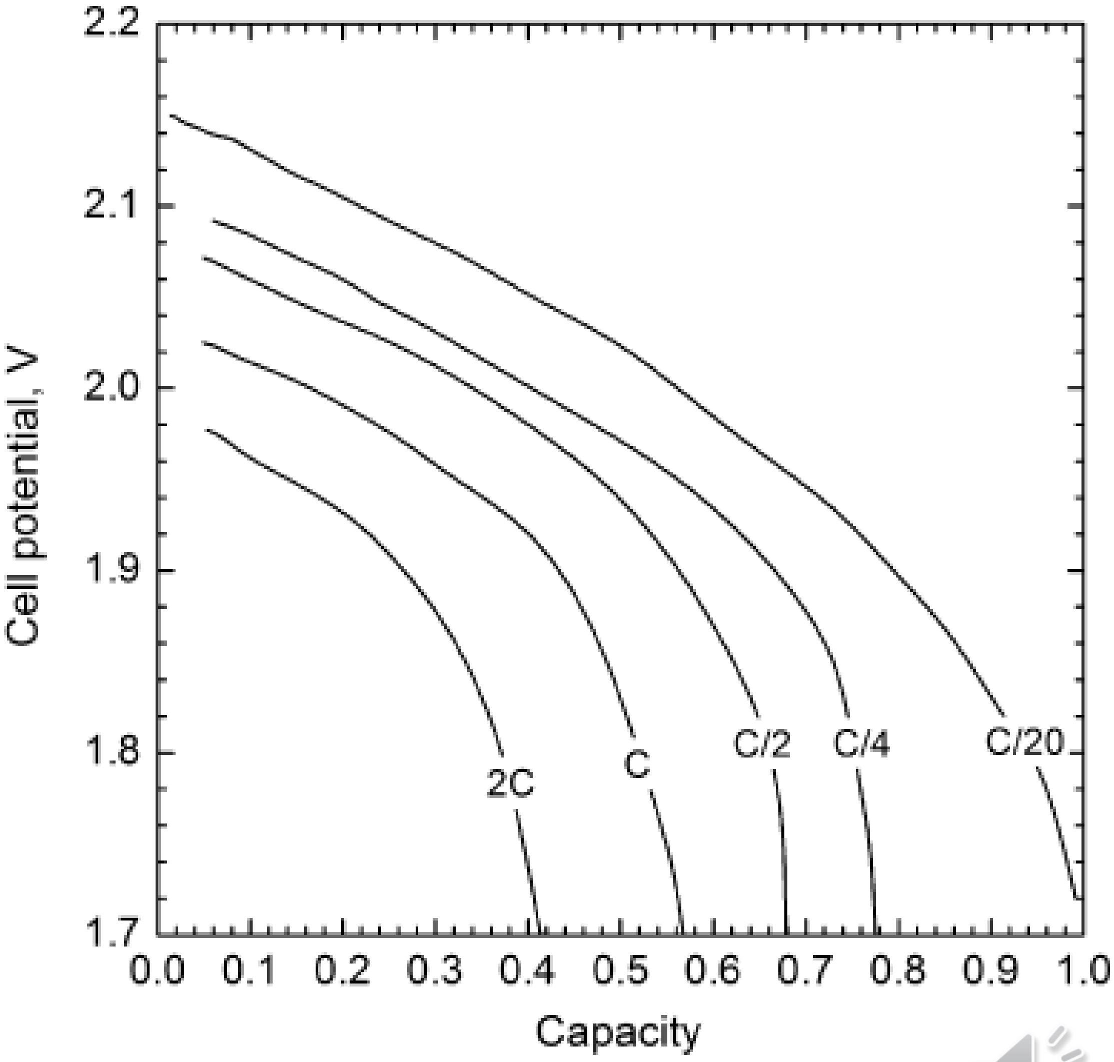
- ❑ **C rate**

  Rate of discharge (or charge):

  $1C$ Current that will take 1 h to fully charge/discharge the electrode (half cell) or the entire battery

  mC Current that take $1/m$ h to fully charge/discharge the single electrode (half cell) or the entire battery

- ❑ **Higher C rates →**
  - Reduced cell potential at the same SOD (or capacity)
  - Lower available capacity at the same cut-off voltage



Lead acid battery

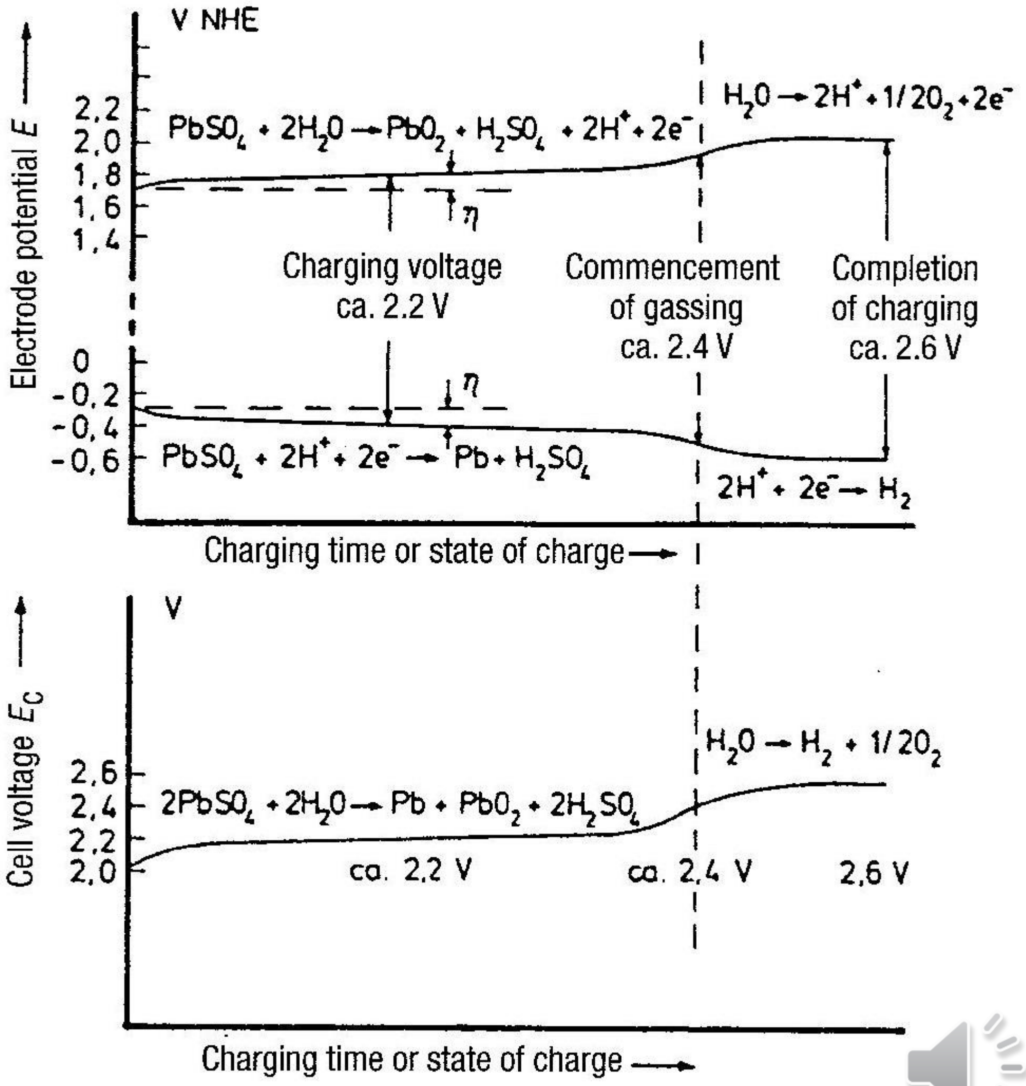
Fuller (2018), Fig. 7.7

Complication of Electrolysis in Charging

- ☐ **For secondary battery, charging voltage has to be larger than equilibrium cell potential**

  To make charging sufficiently fast, the charging voltage may sometimes be higher than the electrochemical window (for electrolyte)

  → may induce side reactions, such as electrolysis of the electrolyte (e.g., water)



Hamann (2007), p. 441

Efficiency of Secondary Battery

- ❑ **Coulombic efficiency**

$$\eta_{coul} = \frac{number\ of\ coulombs\ on\ discharge}{number\ of\ coulombs\ on\ charge} \times 100\%$$

- ❑ **Voltage efficiency**

$$\eta_V = \frac{Average\ cell\ voltage\ on\ discharge}{Average\ cell\ voltage\ on\ charge} \times 100\%$$

- ❑ **Energy efficiency**

$$\eta_{Energy} = \frac{Electrical\ work\ in\ discharge}{Electrical\ energy\ consumed\ in\ charge} \times 100\%$$

$$\eta_{Energy} = \eta_{coul} \cdot \eta_V$$

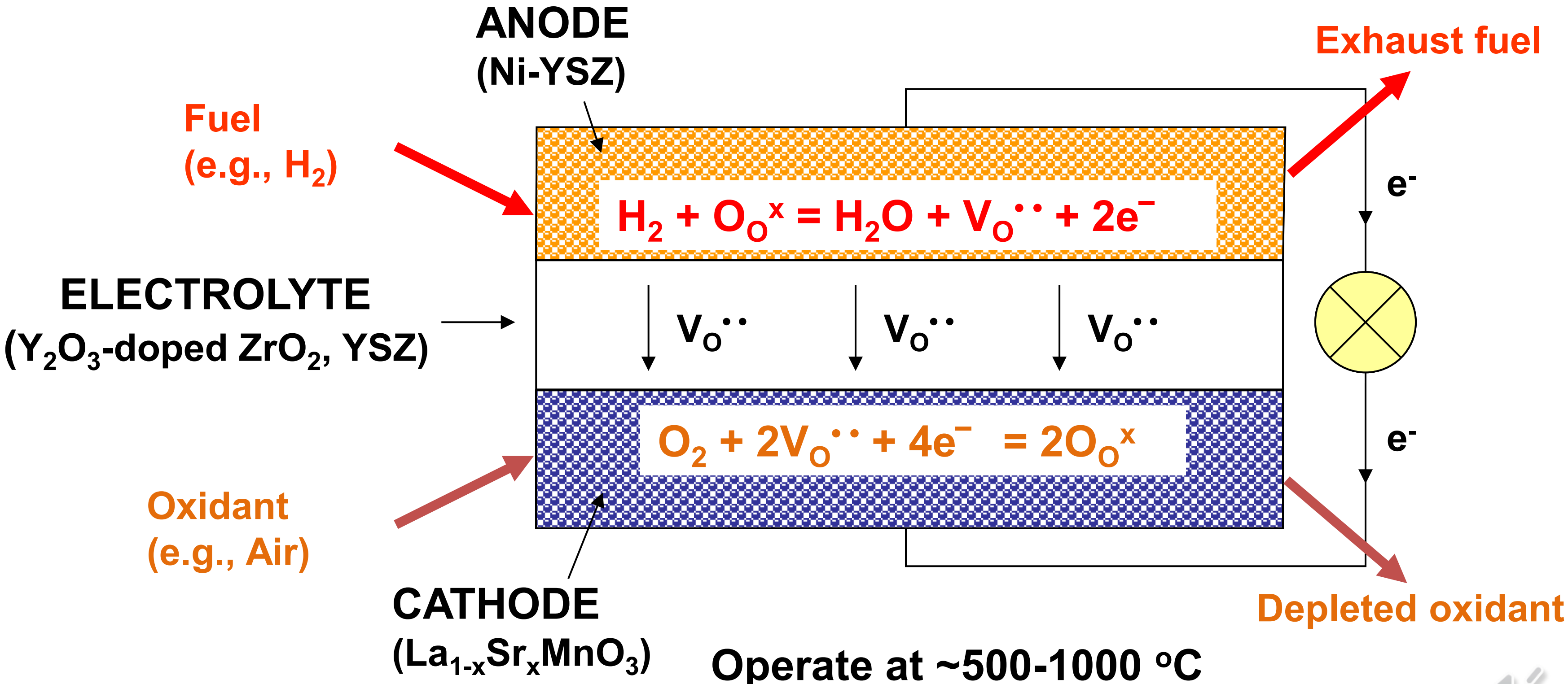
Battery Self-discharge

As the negative electrode is more negative than some electrode reaction related to the electrolyte, side reactions possible, e.g., H_2 evolution in acidic medium, which cause the metal electrode to discharge and lose capacity

Fuel Cell Introduction

- Like battery, but chemical energy is stored in a (gas or liquid) fuel (e.g., H_2, NH_3, methanol), while the oxidant is often O_2

- Example – Solid oxide fuel cell (SOFC)



ANODE
(Ni-YSZ)

Fuel
(e.g., H_2)

$H_2 + O_O^x = H_2O + V_O^{\bullet\bullet} + 2e^-$

Exhaust fuel

e^-

ELECTROLYTE
(Y_2O_3-doped ZrO_2, YSZ)

$V_O^{\bullet\bullet}$ $V_O^{\bullet\bullet}$ $V_O^{\bullet\bullet}$

$O_2 + 2V_O^{\bullet\bullet} + 4e^- = 2O_O^x$

e^-

Oxidant
(e.g., Air)

Depleted oxidant

CATHODE
($La_{1-x}Sr_xMnO_3$)

Operate at ~500-1000 °C

Overall reaction: $2H_2 + O_2 = 2H_2O$

Fuel Cell Classification

Fuel cell type	Ion species	Electrolyte	Anode	Cathode	Operation T & Applications
Direct methanol (DMFC)	H^+	Nafion® polymer electrolyte	Pt-Ru	Pt	20-90°C Portable power
Proton exchange membrane (PEMFC)	H^+	Nafion® polymer electrolyte	Pt	Pt	60-90°C Automobile
Alkaline (AFC)	OH^-	KOH	Ni	Ag	80-100°C
Phosphorous acid (PAFC)	H^+	H_3PO_4	Pt	Pt	~200°C
Molten carbonate (MCFC)	CO_3^{2-}	Li_2CO_3-K_2CO_3	Ni	Li_2O-TiO_2 or Li_2O-NiO compound	~600°C Stationary
Solid oxide (SOFC)	O^{2-}	8 mol% Y_2O_3-doped ZrO_2	Ni	$La_{0.6}Sr_{0.4}Co_{0.2}Fe_{0.8}O_3$	~600-1000°C Stationary

Fuel Cell Open Circuit Voltage (1)

Ambient temperature (25°C) fuel cells

Ion	Anode half cell	Cathode half cell	Full cell
H^+	$2H^+ + 2e^- = H_2$ $E^o(H^+/H_2) = 0$ V vs. SHE	$O_2 + 4H^+ + 4e^- = 2H_2O$ $E^o(O_2/H_2O) = 1.229$ V vs. SHE	$2H_2 + O_2 = 2H_2O$ $E^o = E^o_{cat} - E^o_{an} = 1.229$ V
OH^-	$2H_2O + 2e^- = H_2 + 2OH^-$ $E^o(H_2O/H_2) = -0.828$ V vs. SHE	$O_2 + 4e^- + 2H_2O = 4OH^-$ $E^o(O_2/OH^-) = 0.401$ V vs. SHE	$2H_2 + O_2 = 2H_2O$ $E^o = E^o_{cat} - E^o_{an} = 1.229$ V

Note: both for liquid water

If water is vapor, need to consider $H_2O_{(l)} = H_2O_{(g)}$

Additional free energy change:

$$\Delta G^o_{L\to g} = -\frac{228.582kJ}{mol} - \left(-\frac{237.141kJ}{mol}\right) = 8.559\ kJ/mol$$

Therefore, $E^o_{FC\ water\ vapor} = -\frac{\Delta G^o_{liquid\ water} + \Delta G^o_{L\to g}}{nF} = E^o_{FC\ liquid\ water} - \frac{\Delta G^o_{L\to g}}{nF}$

$$= 1.229V - \frac{8.559kJ/mol}{2\times 96485C/mol} = 1.185V$$

Fuel Cell Open Circuit Voltage (2)

High temperature fuel cells: **overall reaction: $2H_2 + O_2 = 2H_2O$**

$$E_{eq} = E^{o} - \frac{RT}{nF}\ln Q = -\frac{\Delta G^{o}}{nF} - \frac{RT}{nF}\ln\frac{\left(a_{H_2O}\right)^2}{\left(a_{H_2}\right)^2 \cdot a_{O_2}}$$

Free energy change: $\Delta G^{o} = 2\Delta G_f^{o}(H_2O) - 2\Delta G_f^{o}(H_2) - \Delta G_f^{o}(O_2) = 2\Delta G_f^{o}(H_2O)$

$$E_{eq} = -\frac{2\Delta G_f^{o}(H_2O)}{4F} - \frac{RT}{4F}\ln\frac{\left(a_{H_2O}\right)^2}{\left(a_{H_2}\right)^2 \cdot a_{O_2}(cathode)}$$

At 1000 K, from JANAF table (https://janaf.nist.gov/tables/H-064.html)

$$\Delta G_f^{o}(H_2O, g) = -192.590\ kJ/mol$$

in fuel of 97% H_2+3% H_2O and air (a_{O_2} = pO_2 /p^o = 0.21)

$$E_{eq} = -\frac{2\times(-192.590kJ/mol)}{4\times 96485C/mol} - \frac{8.314J/(mol\cdot K)\times 1000K}{4\times 96485C/mol}\times\ln\frac{0.03^2}{0.97^2\cdot 0.21}$$

$$= 0.998V + 0.1185V = 1.12\ V$$

$p\mathrm{O_2}$ from Fuel Cell OCV

High temperature fuel cells: overall reaction: $2H_2 + O_2 = 2H_2O$

$$E_{eq} = E^{\mathrm{o}} - RT\ln Q = -\frac{\Delta G^{\mathrm{o}}}{nF} - \frac{RT}{nF}\ln\frac{\left(a_{H_2O}\right)^2}{\left(a_{H_2}\right)^2 \cdot a_{O_2}(cathode)}$$

On the other hand, at equilibrium, $\Delta G_{eq} = \Delta G^{\mathrm{o}} + RT\ln K = 0$

$$E_{eq} = -\frac{\Delta G^{\mathrm{o}}}{4F} - \frac{RT}{4F}\ln\frac{\left(a_{H_2O}\right)^2}{\left(a_{H_2}\right)^2 \cdot a_{O_2}(cathode)} = \frac{RT\ln K}{4F} - \frac{RT}{4F}\ln\frac{\left(a_{H_2O}\right)^2}{\left(a_{H_2}\right)^2 \cdot a_{O_2}(cathode)}$$

$$= \frac{RT}{4F}\ln\left[K \cdot \frac{\left(a_{H_2}\right)^2 \cdot a_{O_2}(cathode)}{\left(a_{H_2O}\right)^2}\right]$$

$$= \frac{RT}{4F}\ln\left[\frac{\left(a_{H_2O}\right)^2}{\left(a_{H_2}\right)^2 \cdot a_{O_2}(anode)} \cdot \frac{\left(a_{H_2}\right)^2 \cdot a_{O_2}(cathode)}{\left(a_{H_2O}\right)^2}\right] = \frac{RT}{4F}\ln\left[\frac{a_{O_2}(cathode)}{a_{O_2}(anode)}\right]$$

$$a_{O_2}(anode) = a_{O_2}(cathode) \cdot \exp\left(-\frac{4FE_{eq}}{RT}\right)$$

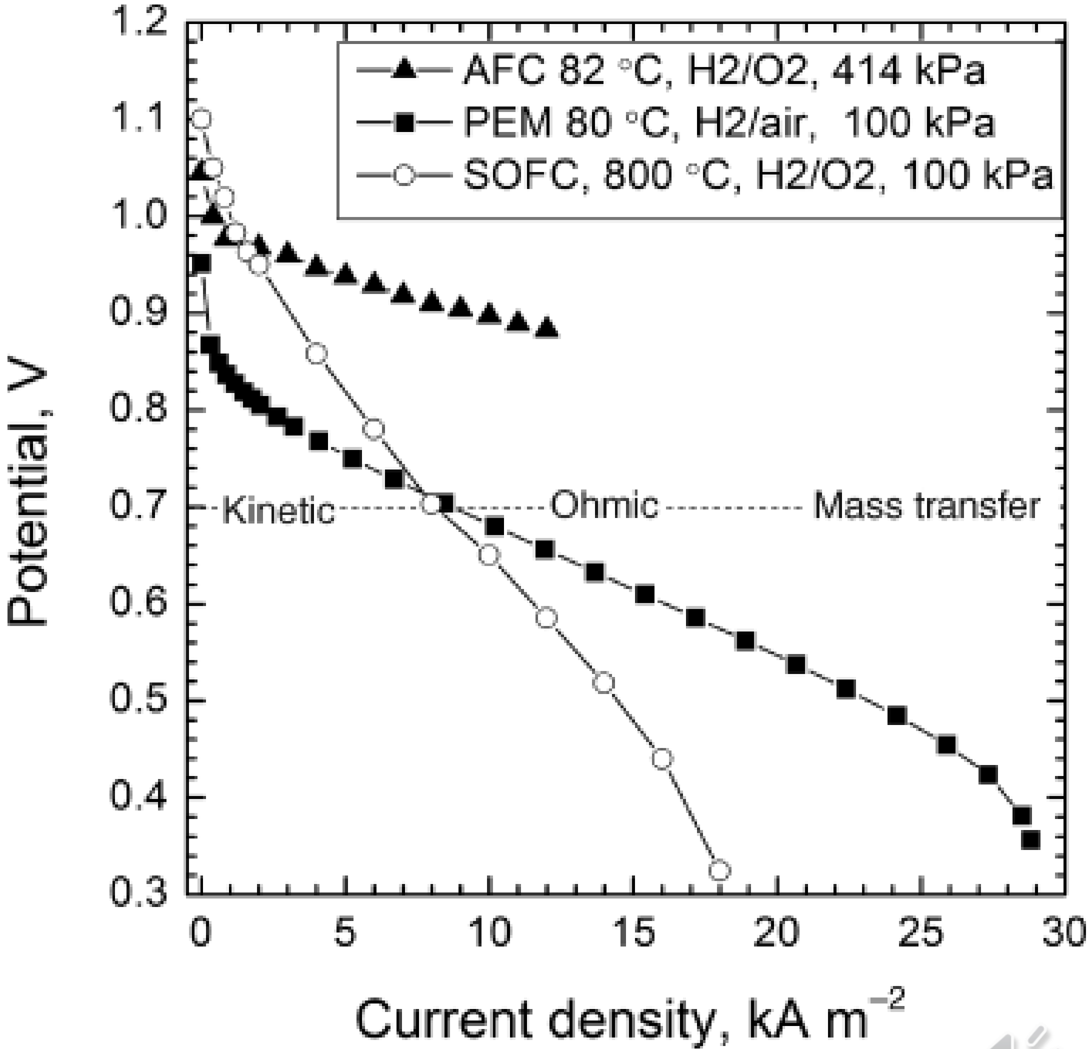
e.g., 1000K, cathode air, E_{eq} = 1.12 V

$$p_{O_2}(anode)/p^{\mathrm{o}} = a_{O_2}(anode) \approx 5 \times 10^{-24}$$

j (or *I*) - *V* Curves for Fuel Cells

Often three regions:

- Low current density – limited by electrode activation/kinetics
- Medium current density – limited by Ohmic resistance (or *IR* loss)
- High current density – limited by mass transfer (e.g., diffusion, convection)

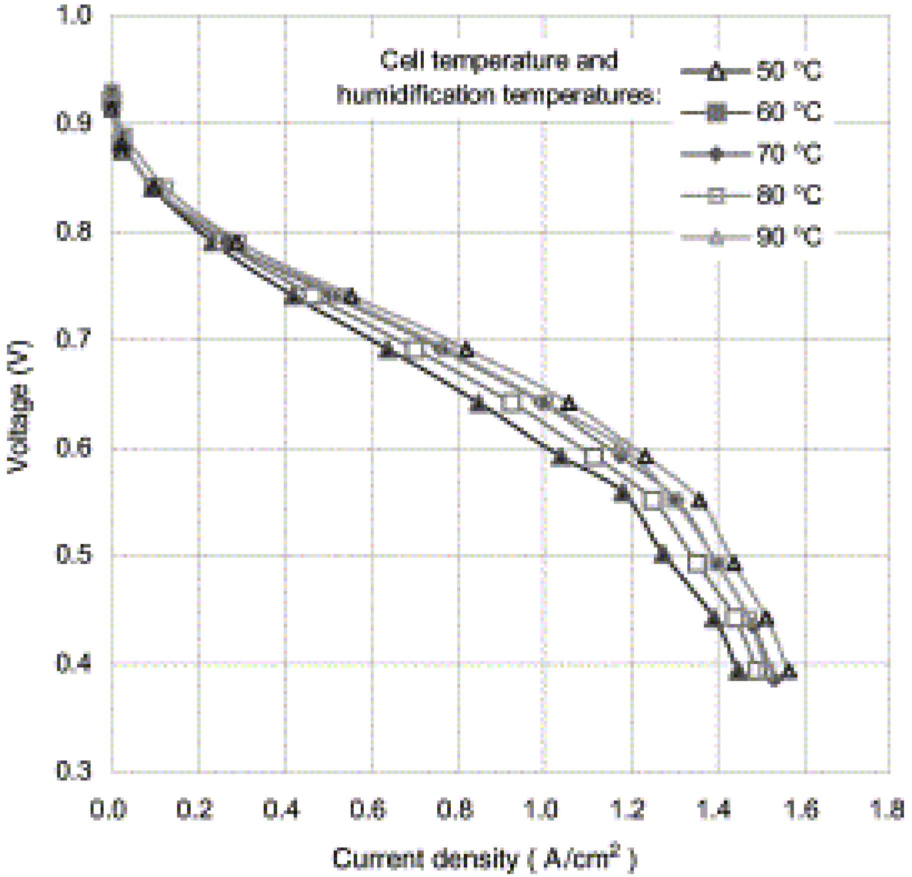


(10 kA/m^2 = 1 A/cm^2)

Fuller (2018), Fig. 9.3

Temperature Effect on Fuel Cell Performance

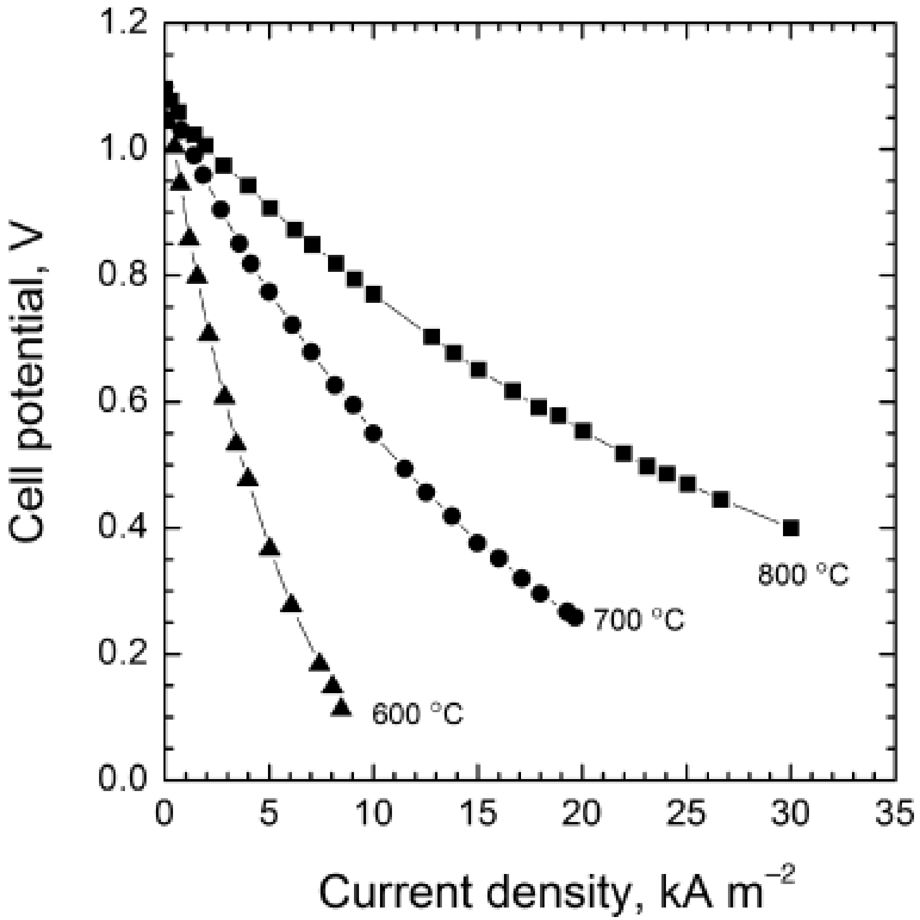
PEMFC



Wang, Inter J Hydrogen Energy (2003), v28(11), 1263

Higher T → **similar** performance

SOFC



Fuller (2018), Fig. 9.19

Higher T → **much higher** performance

Corrosion

❑Definition

- Electrochemical attack of metals due to interactions with oxidizing species in the environment.
- A complete corrosion reaction always involve two parts: oxidation half (cell) reaction and reduction half (cell) reaction

❑Oxidation half-cell reaction

- Metal lose electrons and get oxidized
- Example: $Zn \rightarrow Zn^{2+} + 2e^-$ $E^o(Zn/Zn^{2+})=-0.763V$

❑Reduction half-cell reaction depends on environment:

- If in an acidic environment without oxygen

  $2H^+ + 2e- \rightarrow H_2$ $E^o(H_2/H^+)= 0.0V$

  Overall reaction: $Zn + 2H^+ = Zn^{2+} + H_2$

  Cell potential: $E^o_{cell} = E^o_{cat} – E^o_{an} = 0.0 - (-0.763\ ) V = 0.763 V$
- If In an acidic environment but with oxygen

  $O_2 + 4e^- + 4H^+ \rightarrow 2H_2O$ $E^o(H_2O/O_2)= 1.23 V$

  Overall reaction: $2Zn + 4H^+ +O_2 = 2Zn^{2+} + 2H_2O$

  Cell potential: $E^o_{cell} = E^o_{cat} – E^o_{an} = 1.23 - (-0.763) V = 1.993 V$
- Both energetically favorable, which one dominates depend on environment

Passivation (1)

- **When the corrosion product is a dense solid that have very low solubility and does not readily permit passage of water, ions, and other active species (e.g., O_2 gas), passivation would occur.**

- **Example of Si passivation:**
  - For Si in neutral or acidic aqueous medium, the anodic half-cell reaction can be

  $Si(s) + 2H_2O = SiO_{2(s)} + 4H^+$ $E^o(Si/SiO_2) = -0.91$ V
  - In principle, the cathodic half-cell reaction can be either

  Without oxygen in acidic/neutral water: $2H^+ + 2e^- \rightarrow H_2$ $E^o(H_2/H^+) = 0.0$V

  With oxygen in acidic/neutral water: $O_2 + 4e^- + 4H^+ \rightarrow 2H_2O$ $E^o(H_2O/O_2) = 1.23$ V

  For either one, the electrochemical cell potential will be highly positive, meaning there is a very large driving force for the reactions to proceed
  - However, because the Si corrosion (oxidation) product of SiO_2 is very dense and impermeable to the typical corrosive species including H_2O, H^+, or O_2, despite the high cell potential, the full cell reactions do not really proceed at detectable rate in neutral or acidic condition

Passivation (2)

❑Example of Si passivation:

- On the other hand, in basic medium, the anodic half-cell reaction can be:

$Si + 6OH^- = SiO_3^{2-} + 3H_2O + 4e^-$ $E^o(Si/SiO_3^{2-}) = -1.697$ V

- The cathodic half-cell reaction can be:

$O_2 + 2H_2O + 4e^- = 4OH^-$ $E^o(OH^-/O_2) = 0.401$ V

- The overall reaction will be:

$Si + O_2 + 2OH^- = SiO_3^{2-} + H_2O$

$E^o_{cell} = E^o_{cat} - E^o_{an} = 0.401 - (-1.697)$ V $= 2.098$ V

- Such a reaction could happen in a strongly basic condition (sometimes with heat necessary)

❑Similar cases for Al in neutral water or Fe in concentrated nitric acid

Pourbaix (E-pH) Diagram for Aluminum

❑Boundary lines

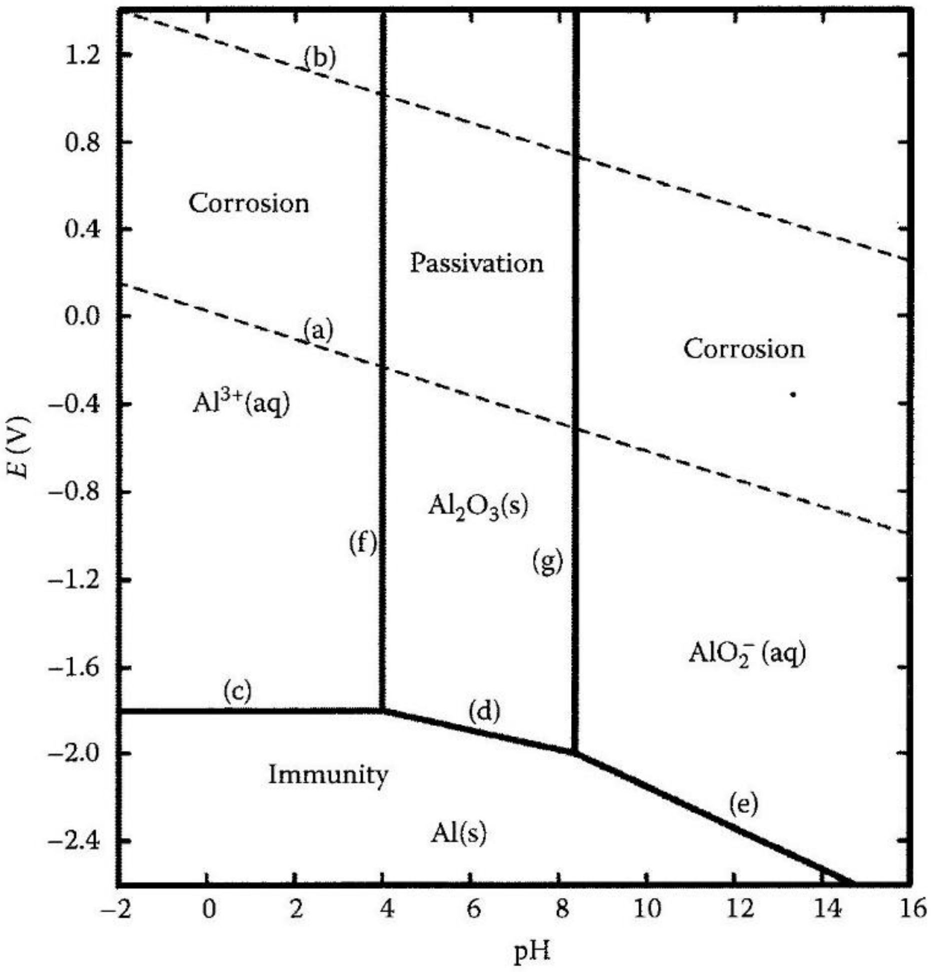
(a) $2H^+ + 2e^- = H_2$
(b) $0.5O_2 + 2H^+ + 2e^- = H_2O$
(c) $Al^{3+} + 3e^- = Al$
(d) $0.5Al_2O_3 + 3H^+ + 3e^- = Al + 1.5H_2O$
(e) $AlO_2^- + 4H^+ + 3e^- = Al + 2H_2O$
(f) $2Al^{3+} + 3H_2O = Al_2O_3 + 6H^+$
(g) $Al_2O_3 + H_2O = 2AlO_2^- + 2H^+$

❑Regions

Immunity: low enough potential that it is protected thermodynamically
Corrosion: in strong acid (pH<~4)
Corrosion: in strong base (pH > ~8)
Passivation: around neutral

❑Limitations of Pourbaix diagrams

- No information about corrosion rate



Polarization Curve for Metal Corrosion

- **Polarization curve for metal corrosion**

  Change the potential of the metal (with respect to a reference electrode such as SHE or Ag/AgCl) and measure the change in current

- **Example**

  Polarization curve for a metal Me in an acidic solution w/o dissolved O_2:

  Region 0: $E < E(M^{z+}/M)$,

  Essentially no current, no corrosion

  Region 1: $E(M^{z+}/M) < E < E_F$,

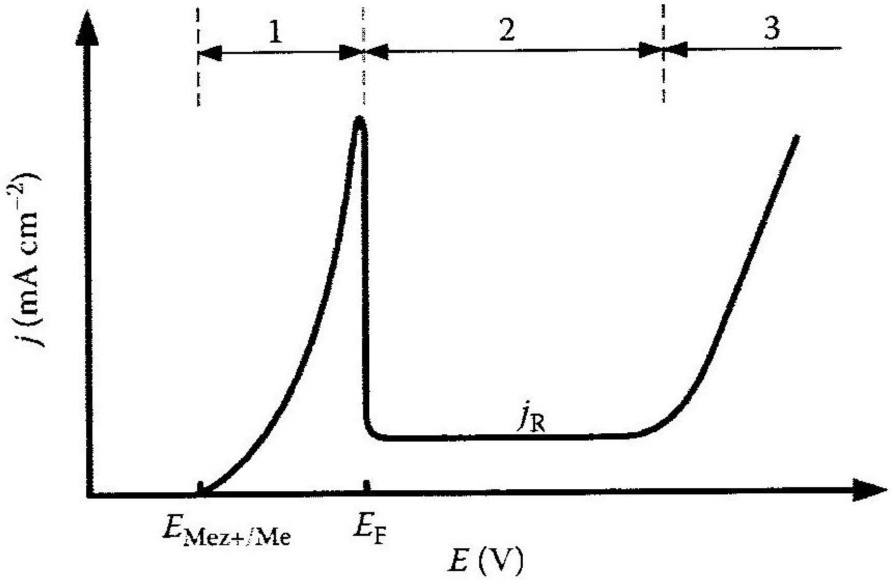
  Rapid increase in corrosion rate with more positive E

  Region 2: $E_F < E < E_t$,

  Low current due to passivation

  Region 3: E > Et

  Rapid increase in current due to either break down of passivation film or alternative reaction (e.g., water electrolysis)



Lvov (2015), p. 178

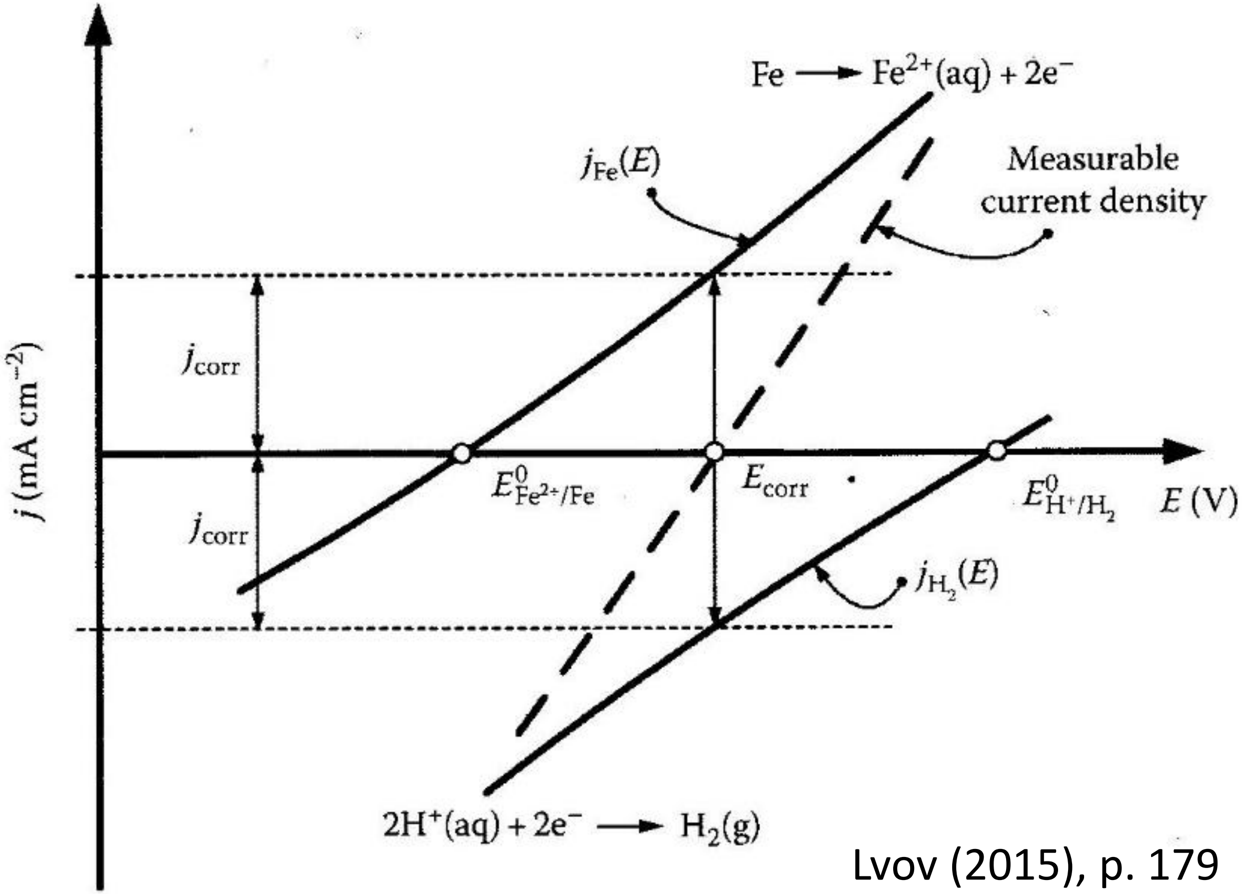
Corrosion Potential & Current Density

E_{corr} & j_{corr}

Using Fe corrosion in dilute non-oxidizing acid solution without dissolved oxygen as an example, two separate polarization (*I* vs. *E*) curves:

- one for the redox of the metal (Fe)
- one for the redox of the other reducible species in environment:

When the absolute value for the anodic current for the metal of interest (Fe in this case) equals the cathodic current for the environment species, a stable corrosion potential E_{corr} and current density j_{corr} are established



Lvov (2015), p. 179

Evans Diagram for Corrosion

Another way

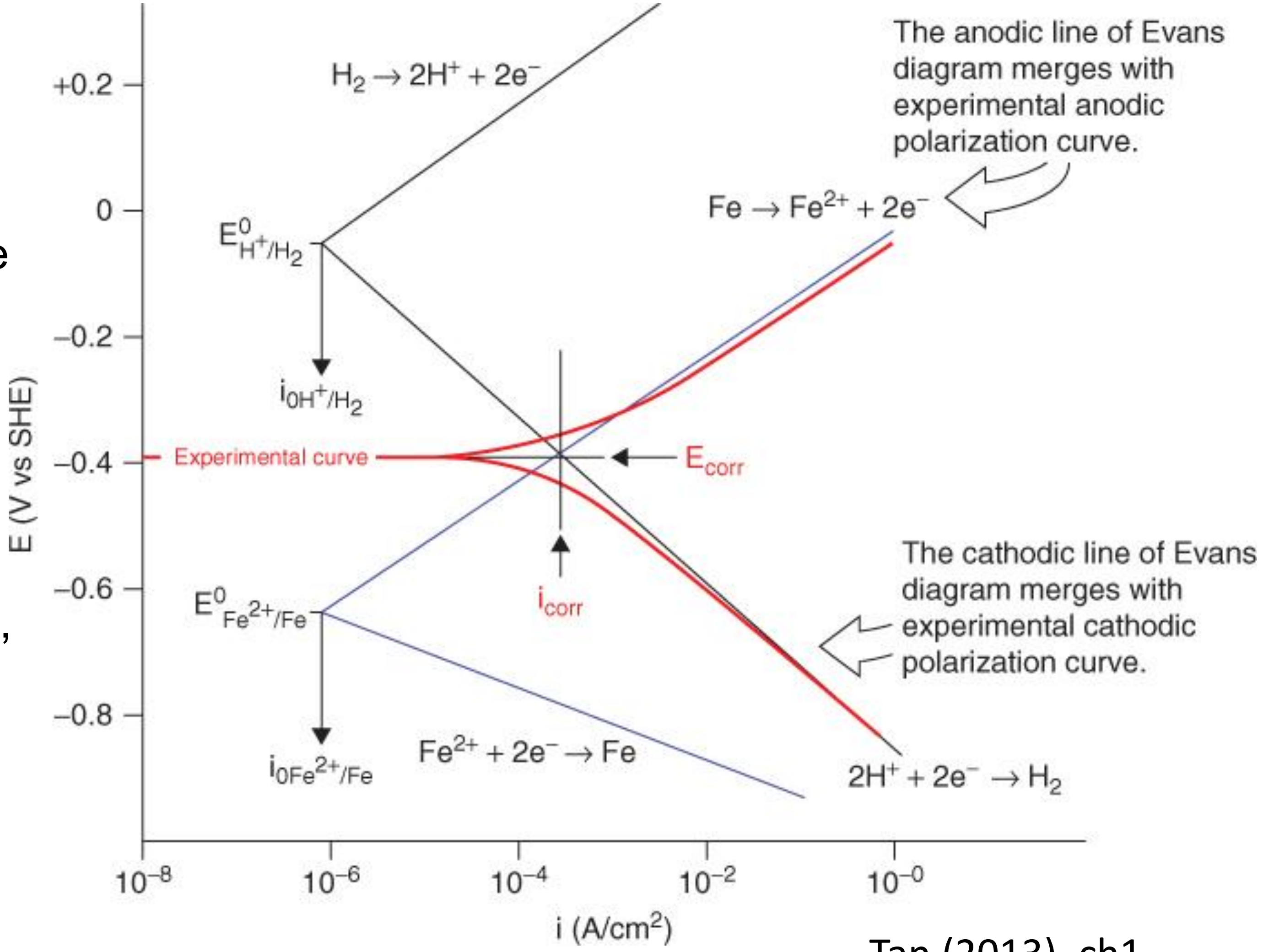
Use Fe corrosion in dilute acid without dissolved oxygen

Only plot absolute value of current density, two sets of curves can be obtained:

One for Fe redox

One for H_2 redox

Where they cross, gives E_{corr} and j_{corr}



The anodic line of Evans diagram merges with experimental anodic polarization curve.

The cathodic line of Evans diagram merges with experimental cathodic polarization curve.

Tan (2013), ch1

Corrosion Rate & Prevention

❑Corrosion rate (CR)

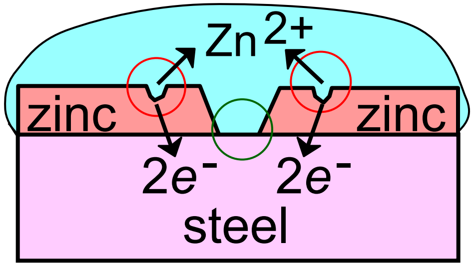
Often given as depth of metal corroded per unit time:

$$CR = \frac{j_{corr} M_{Me}}{\rho_{Me} z F}$$

❑Corrosion prevention

- Replacing materials susceptible to corrosion by those that are not susceptible to environment attack
  - Replace steel with plastics or precious metals
- Preventing the contact between active metal and its environment
  - External coating via painting
  - (Self) passivation: Al, Ti, or stainless steel
- Cathodic protection:
  - To protect Fe, connect Fe with Zn (e.g., via Zn plating) as $E^o(Zn^{2+}/Zn) = -0.762V$ is more negative than $E^o(Fe^{2+}/Fe) = -0.447V$, Fe will not corrode until Zn is consumed. This way, Zn is called sacrificial anode

Galvanized Steel



Adapted from Fig. 17.23, *Callister & Rethwisch 8e.*

e.g., zinc-coated nail

Homework

- ☐ **Raise THREE (3) question that you don't understand for lecture videos**

  In case you have understood everything and don’t have that many questions, please give corresponding number of multiple-choice problem (together with your answer) that you feel can be used to check a student's understanding.

  An example multiple-choice problem could be:

  Which of the units below can be the unit for current density j?

  a) A
  b) A/cm^2 (Answer)
  c) V
  d) C