IIntroduction & Overview of Electrode Processes (Ch. 1)

Introduction
Electrochemical cells & reactions: thermodynamics and potentials

Nonfaradaic processes & electrode-solution interface
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Reaction rates and current
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Introduction

Electrochemistry: passage of electric current → chemical changes
chemical reactions → production of electric energy

Electrochemical cells & reactions

Electrode: junction between electronic conductor and ionic conductor (electrolyte) that the chemistry of electrochemistry occurs

e.g., solid metal, liquid metal, carbon, semiconductor

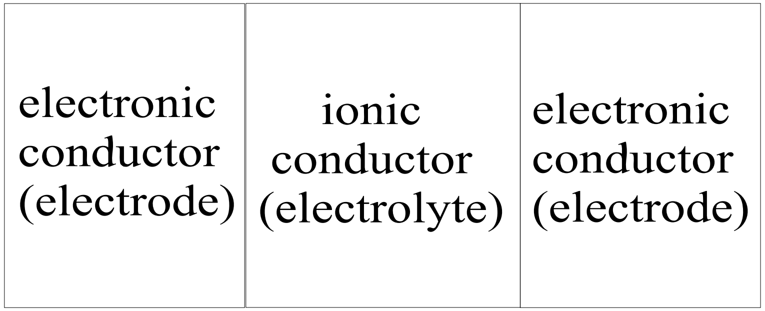
Electrolyte: substance that produces ions so enhance the electrical conductivity

e.g., liquid(aqueous, non-aq), gas(NH_3), solid electrolyte (polymer, ceramic)

Electrochemical cell: two electrodes + electrolyte

an ionic conductor sandwiched between two electronic conductors

e.g., aqueous solution of electrolyte between two pieces of metal, solid electrolyte between two metals



Cell potential (V, volts): 1 V = 1 J/C, energy to drive charge between electrodes

Electrochemical cell notation:

$Zn/Zn^{2+}, Cl^{-}/AgCl/Ag$ $Pt/H_2/H^{+}, Cl^{-}/AgCl/Ag$

slash(/): phase boundary,

comma(,): two components in the same phase

double slash(//): phase boundary with negligible phase boundary

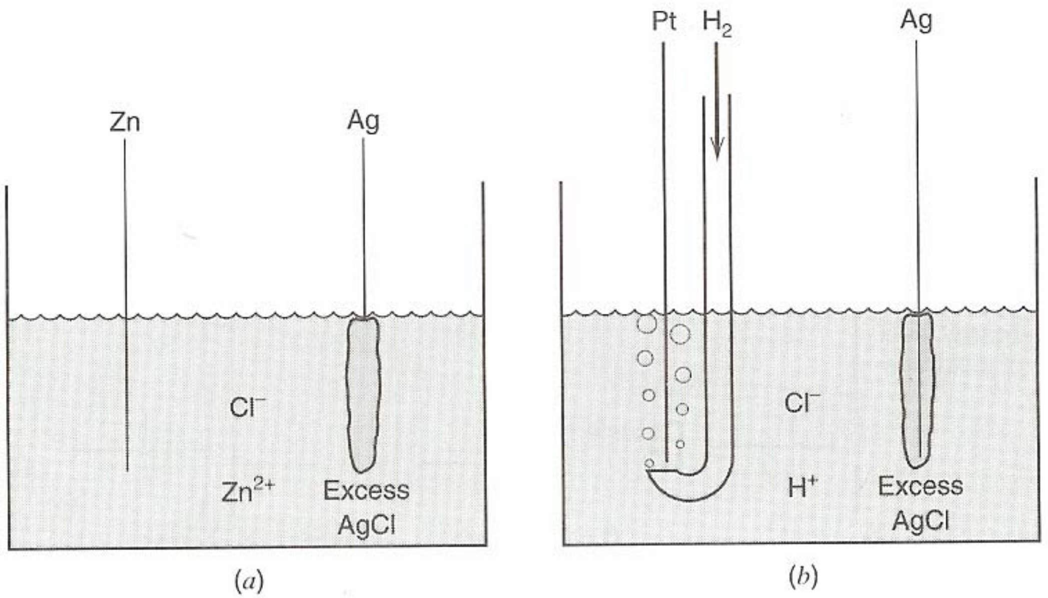


Figure 1.1.1 Typical electrochemical cells. (*a*) Zn metal and Ag wire covered with AgCl immersed in a $ZnCl_2$ solution. (*b*) Pt wire in a stream of H_2 and Ag wire covered with AgCl in HCl solution.

Two half-reactions

(i) interest in two equal electrodes (e.g., battery)

(ii) interest in **one electrode** only

- Working (or indicator) electrode(WE): electrode of interest
- Reference electrode(RE): second electrode, measure potential of WE with respect to RE
- Electrode potential $E = E_{work} – E_{ref}$

Reference electrodes

- SHE (standard hydrogen electrode) or NHE(normal hydrogen electrode): universally accepted standard: $Pt/H_2(a=1)/H^+(a=1, aqueous)$

$$H^+(aq, a=1) + e^- = 1/2H_2(g, 10^5\ Pa) \quad E = 0\ V$$

- SCE (saturated calomel electrode): $Hg/Hg_2Cl_2/KCl(sat)$

$$Hg_2Cl_2(s) + 2e^- = 2Hg + Cl^- \quad E_{ref} = 0.242\ V\ vs.\ NHE$$

- Ag/AgCl (silver-silver chloride): Ag/AgCl/KCl(sat or OO M)

$$AgCl(s) + e^- = Ag(s) + Cl^-(aq) \quad E_{ref} = 0.197\ V\ \text{with saturated KCl}$$

Potentials of reference electrodes

E(RHE) = E(NHE) + 0.05916pH

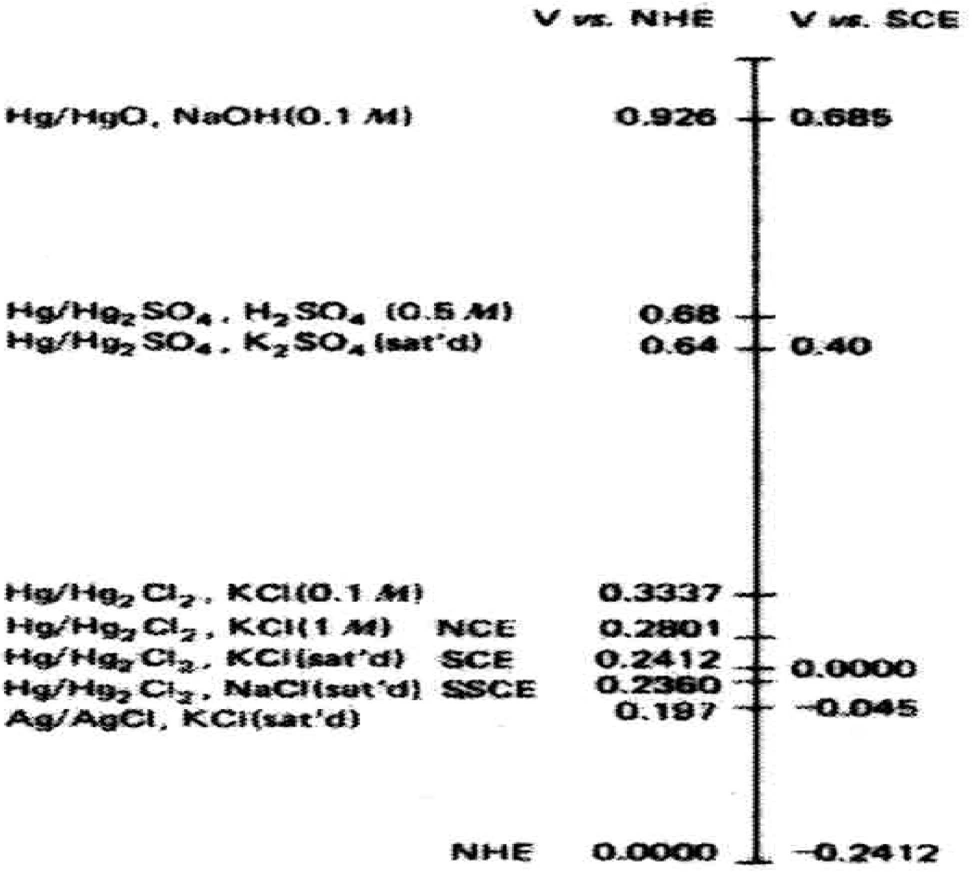
E(SCE) = E(NHE) – 0.242

E(Ag/AgCl) = E(NHE) – 0.2223

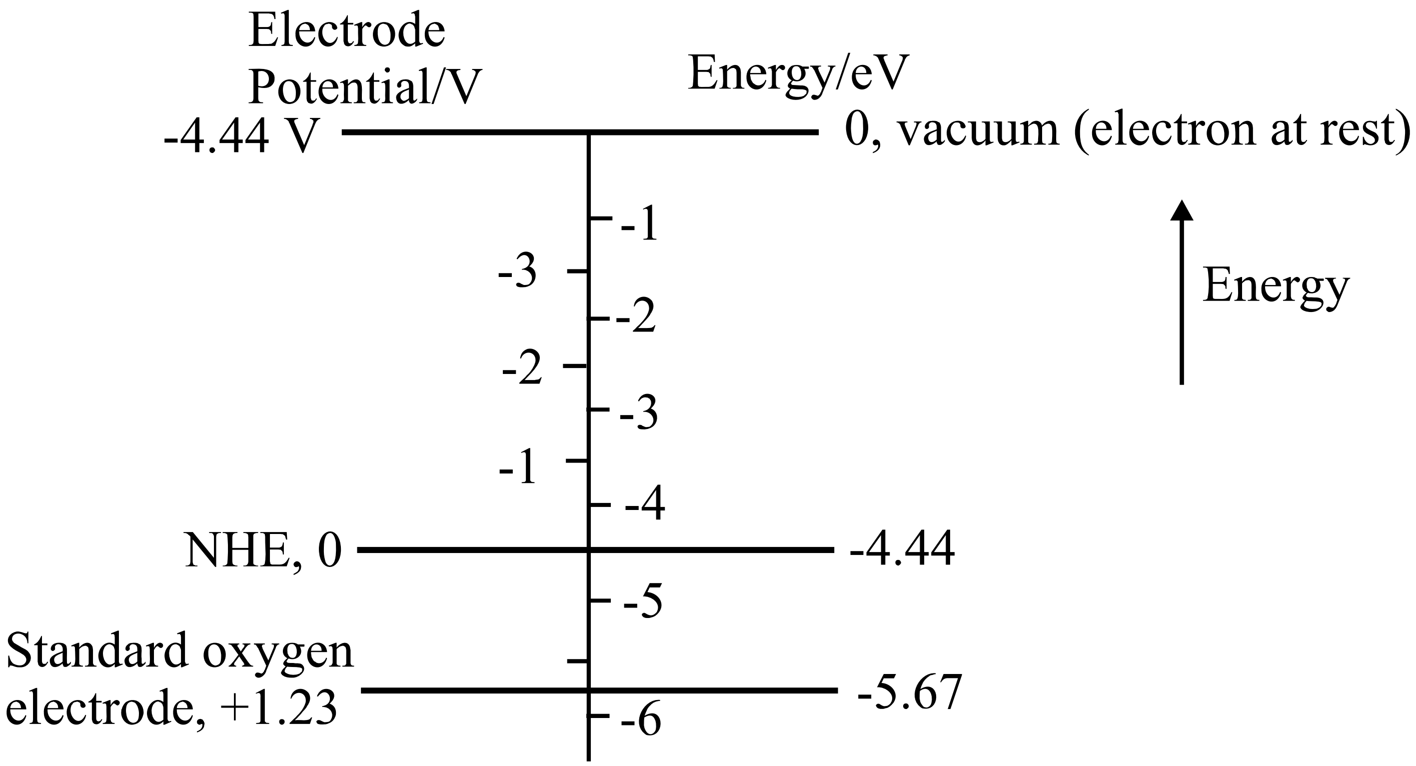
E(Ag/AgCl, sat.KCl) = E(NHE) – 0.196

E(Hg/HgO 1M KOH) = E(NHE) – 0.1100 + 0.05946pH

E(Hg/Hg_2SO_4) = E(NHE) – 0.6152



Potential vs. energy (vs. vacuum)



Electrode Potential/V

Energy/eV

-4.44 V — 0, vacuum (electron at rest)

-3

-2

-1

NHE, 0 — -4.44

Standard oxygen electrode, +1.23 — -5.67

-1

-2

-3

-4

-5

-6

Energy

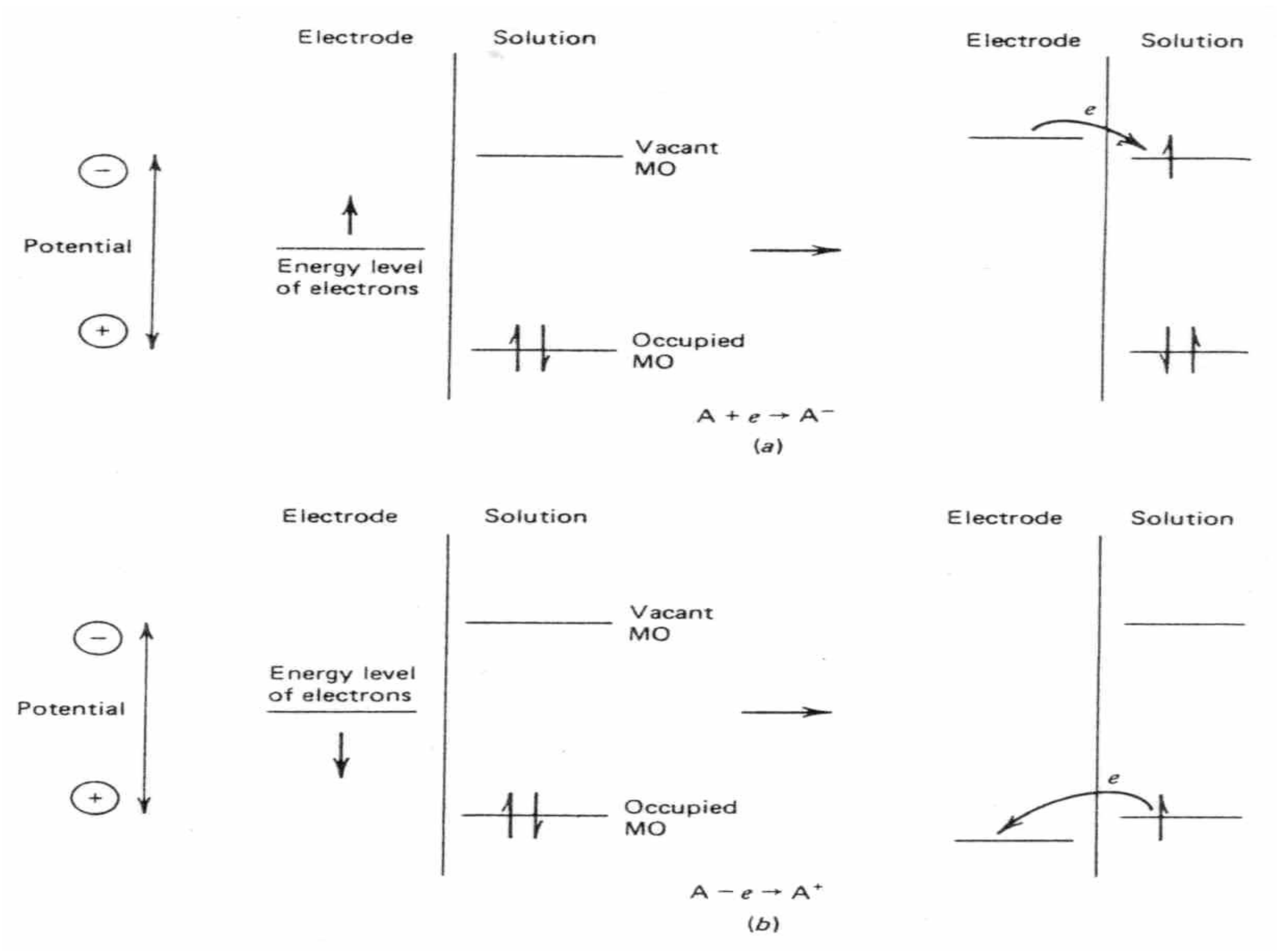
Controlling potential of the working electrode with respect to the reference → controlling the energy of the electrons within the working electrode

More negitive potential → energy of electrons is raised → reach a level to occupy vacant states (LUMO) on species in the electrolyte → flow of electrons from electrode to solution (reduction current)

More positive potential → electron flow from solution (HOMO) to electrode (oxidation current)

Standard potential (E^0): critical potential at which these processes occur for the specific chemical substances in the system (p. 808, standard potential)

Applying potential from its equilibrium (or its zero-current)



Electrochemical experiment:

- Electrochemical cell (WE/electrolyte/RE) + power supply
- Faraday's law: charge(Q, C(1 C = 6.24 x 10^{18} e^-) vs. extent of chemical reaction
  "the passage of 96485.4 C causes 1 equivalent of reaction (e.g., consumption of 1 mole of reactant or production of 1 mole of product in a one-electron rxn)"

$$F = N_A Q_e = (6.02 \times 10^{23}\ \text{mol}^{-1})(1.6022 \times 10^{-19}\ \text{C}) = 96485\ \text{Cmol}^{-1}$$

- Current (i): rate of flow of coulombs (or electrons) (1 A = 1 C/s)

current-potential (i vs. E) curve

open-circuit potential (zero-current potential or rest potential): potential measured by a high impedance voltameter. Since no current flows, it makes no difference if the circuit is interrupted, as by opening the switch

(i) redox couple (equilibrium established, e.g., Fig. 1.1.1): calculation of open-circuit potential is possible from the standard potentials of the half-reactions via the Nernst equation (open-circuit potential = equilibrium potential)

(ii) Overall equilibrium can not be established

e.g., Pt/H^+, Br^-/AgBr/Ag

Ag/AgBr electrode: AgBr + e = Ag + Br^-, 0.07 V

Pt/H^+, Br^-: not at equilibrium

→ equilibrium potential does not exist

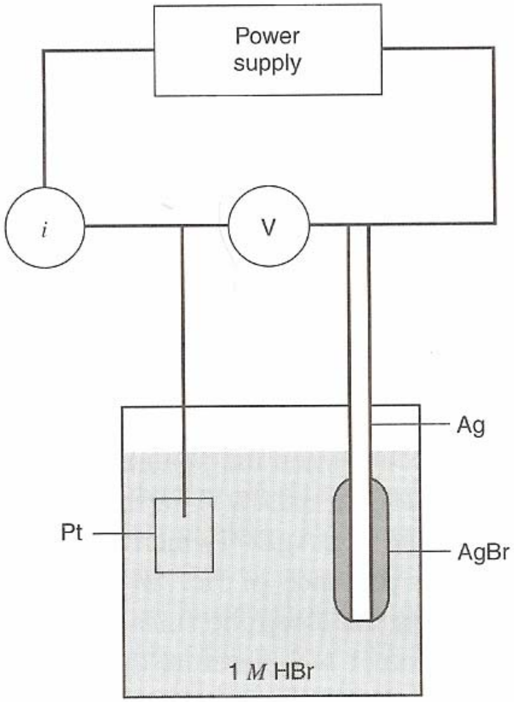


Figure 1.1.3 Schematic diagram of the electrochemical cell Pt/HBr(1 *M*)/AgBr/Ag attached to power supply and meters for obtaining a current-potential (*i*-*E*) curve.

OCV is not available from calculation

Then, open-circuit potential for this case? i-E curve?

1st electrode rxn: $2H^+ + e \rightarrow H_2$, -0.07 V vs. Ag/AgBr

2nd electrode rxn: oxidation of Br^- to Br_2

$Br_2 + 2e = 2Br^-$, 1.09 V vs. NHE (= 1.02 V vs. Ag/AgBr)

- Background limits (+1.02 V to +0.07 V)
- OCV: somewhere in background limits

(experimentally found, depends on impurities)

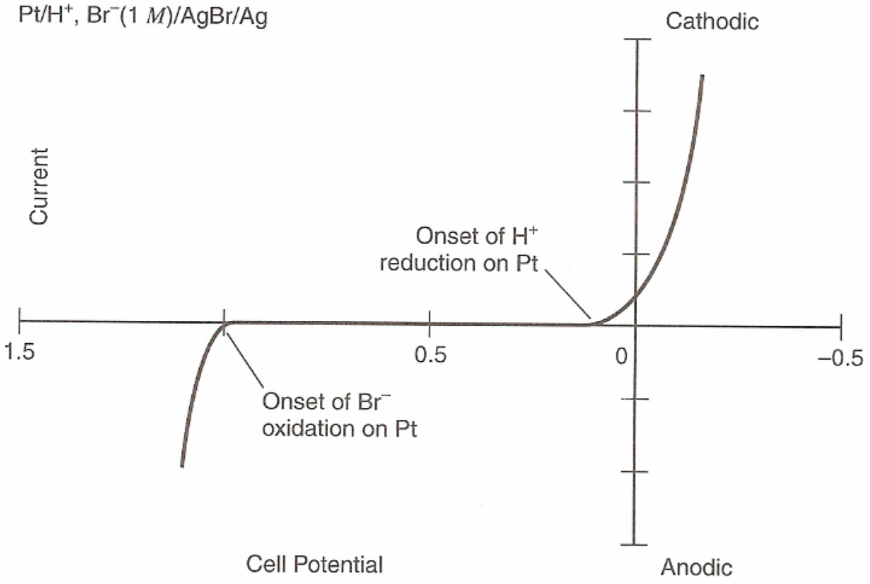


Figure 1.1.4 Schematic current-potential curve for the cell Pt/H$^+$, Br$^-$(1 *M*)/AgBr/Ag, showing the limiting proton reduction and bromide oxidation processes. The cell potential is given for the Pt electrode with respect to the Ag electrode, so it is equivalent to E_{Pt} (V *vs.* AgBr). Since $E_{Ag/AgBr} =$ 0.07 V *vs.* NHE, the potential axis could be converted to E_{Pt} (V *vs.* NHE) by adding 0.07 V to each value of potential.

e.g., Hg/H^+, $Br^-/AgBr/Ag$
OCV is not available from calculation
$2H^+ + e \rightarrow H_2$, 0.0 V vs. NHE (thermodynamic), slow rate: much more negative E
(“overpotential”)
oxidation of Hg to Hg_2Br_2 at 0.14 V
- Background limits (-0.9 V to +0.0 V): depends on electrode material & solution

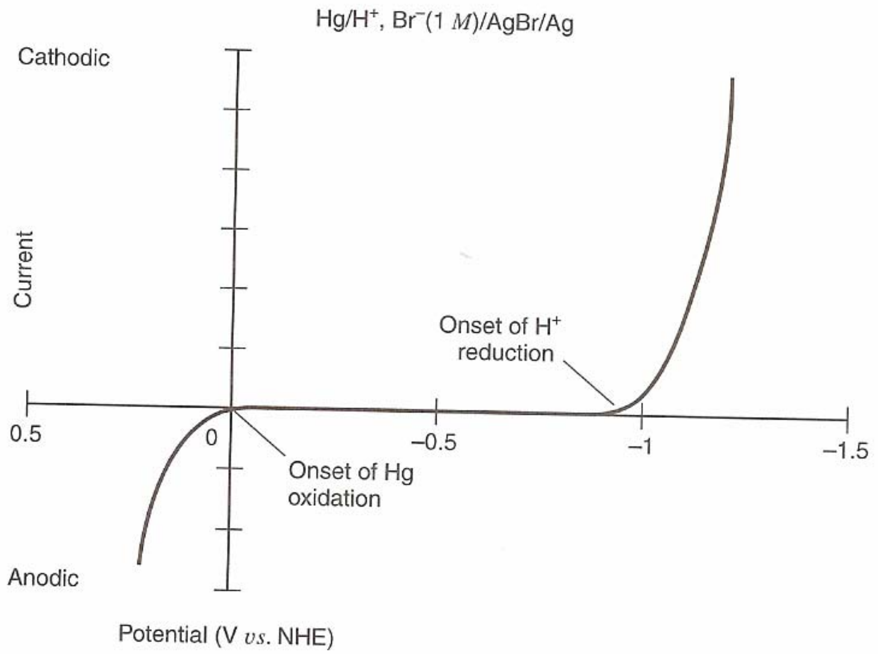


Figure 1.1.5 Schematic current-potential curve for the Hg electrode in the cell Hg/H^+, Br^-(1 *M*)/AgBr/Ag, showing the limiting processes: proton reduction with a large negative overpotential and mercury oxidation. The potential axis is defined through the process outlined in the caption to Figure 1.1.4.

e.g., Hg/H^+, Br^-, Cd^{2+}(10^{-3} M)/AgBr/Ag
$Cd^{2+} + 2e = Cd(Hg)$ -0.4 V vs. NHE

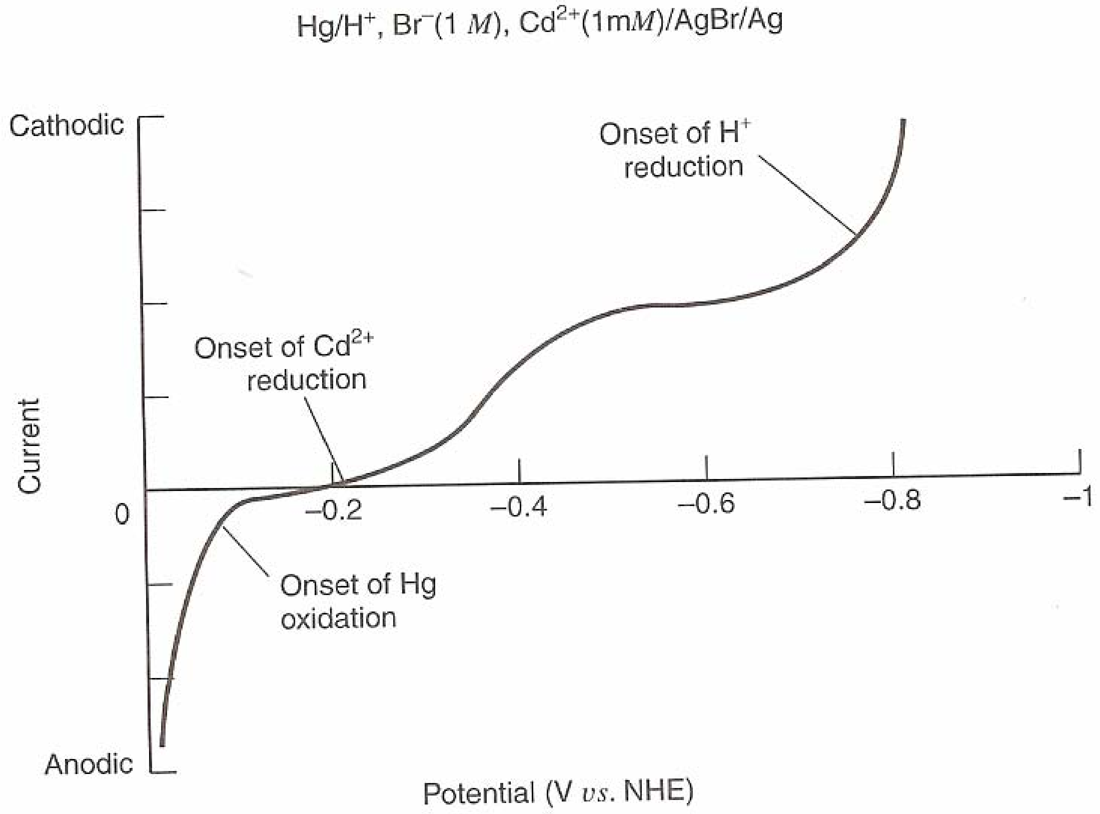
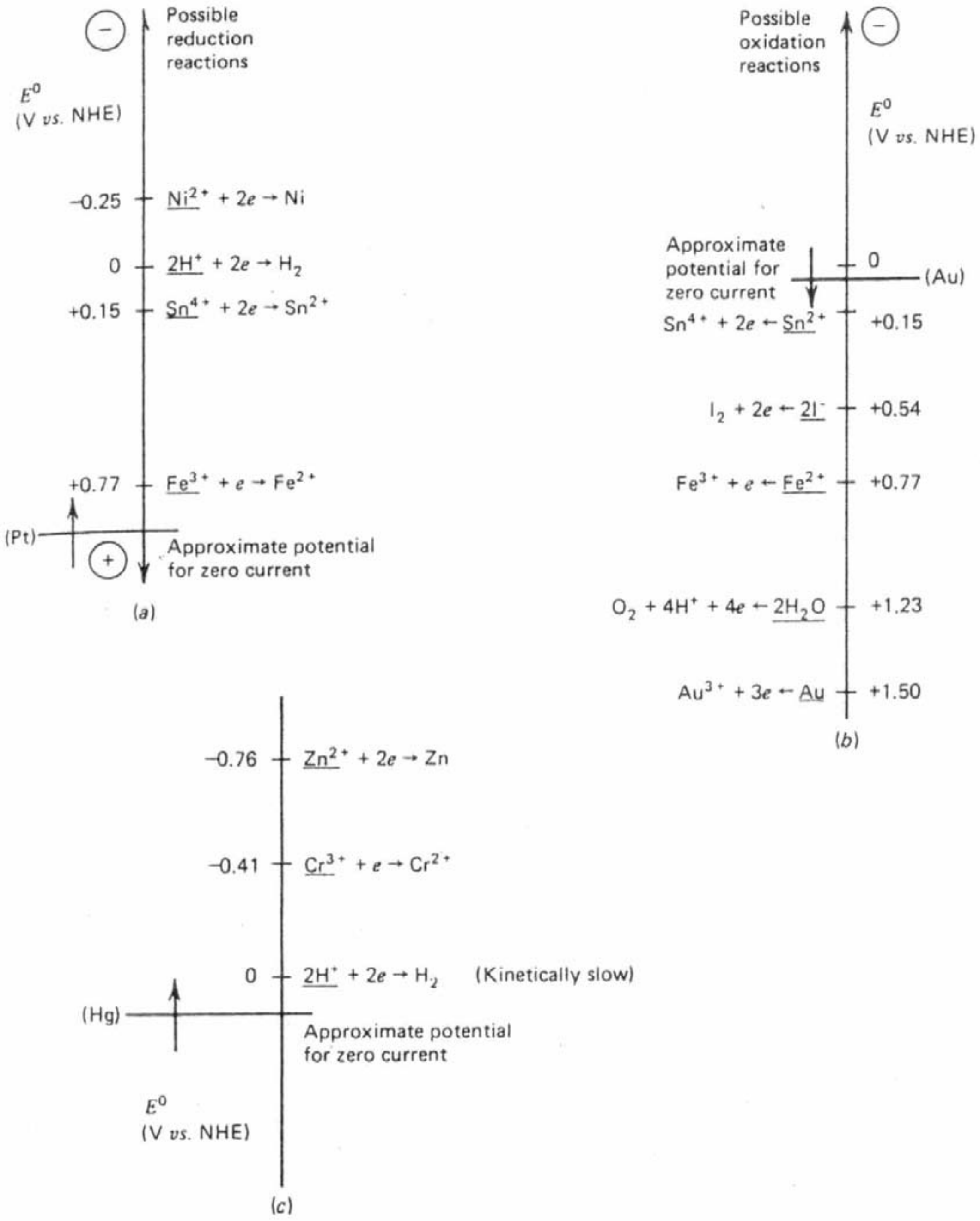


Figure 1.1.6 Schematic current-potential curve for the Hg electrode in the cell Hg/H^+, Br^-(1 *M*),Cd^{2+}(10^{-3} M)/AgBr/Ag, showing reduction wave for Cd^{2+}.

- Potential moved from OCV toward more negative potential: reduced more positive E^0 first
- Potential moved from OCV toward more positive potential: oxidized more negative E^0 first
- consider slow kinetics: slow hydrogen evolution in Hg → Cr^{3+} reduction first in Figure (c)



Faradaic and nonfaradaic processes

Faradaic process: charges (e.g., electrons) are transferrred across the electrode-electrolyte interface. Electron transfer causes oxidation and reduction to occur: governed by Faraday's law (the amount of chemical reaction caused by the flow of current is proportional to the amount of electricity passed) → "charge transfer electrode"

Nonfaradaic process: no charge transfer reactions occur because of thermodynamically and kinetically unfavorable (0 to 0.8 V in Fig.1.1.5). Adsorption/desorption can occur. The structure of the electrode-solution interface can change with changing potential or solution composition. Transiently external current can flow.

Both faradaic and nonfaradaic processes occur when electrode reactions take place

Nonfaradaic processes and the electrode-solution interface

Ideal polarized electrode (ideal polarizable electrode) (IPE): no charge transfer over limited potential ranges

e.g., Hg in KCl (+0.25 V to -2.1 V), alkane thiol on Au

Capacitance and charge of an electrode

- IPE interface = capacitor (two metal sheets separated by a dielectric material)

$$q/E = C$$

q: charge stored on the capacitor (C, coulomb)

E: potential across the capacitor (V), C: capacitance (F, farad)

- During this charging process, a current ("charging current") will flow
- 2 V battery across 10 μ F capacitor

→ current will flow until 20 μ C accumulated

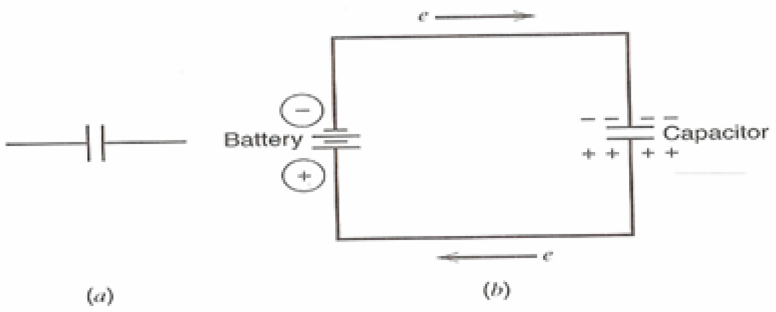


Figure 1.2.1 (*a*) A capacitor. (*b*) Charging a capacitor with a battery.

Electrode-solution interface: $q^M = -q^S$ “electrical double layer”

q^M: very thin layer (<0.1 Å), charge density (μ C/cm^2) $\sigma^M = q^M/A$

Typical double-layer capacitance (C_d): 10 ~ 40 μ F/cm^2

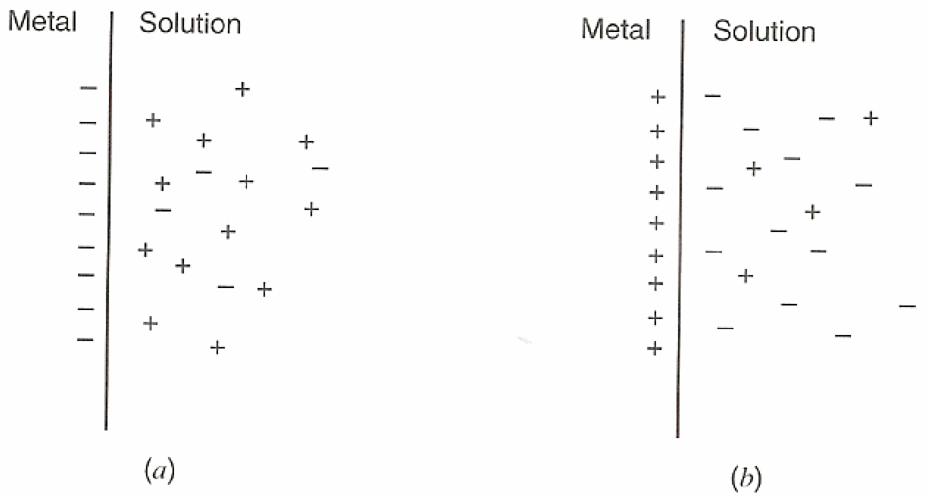


Figure 1.2.2 The metal–solution interface as a capacitor with a charge on the metal, q^M, (*a*) negative and (*b*) positive.

Electrical double layer

Double layer: several layers

Inner layer (compact, Helmholtz, Stern):

solvent, specifically adsorbed species

Outer Helmholtz plane (OHP): solvated ions

Diffusion layer: extends from OHP to the bulk

(~100 Å in >10^{-2} M)

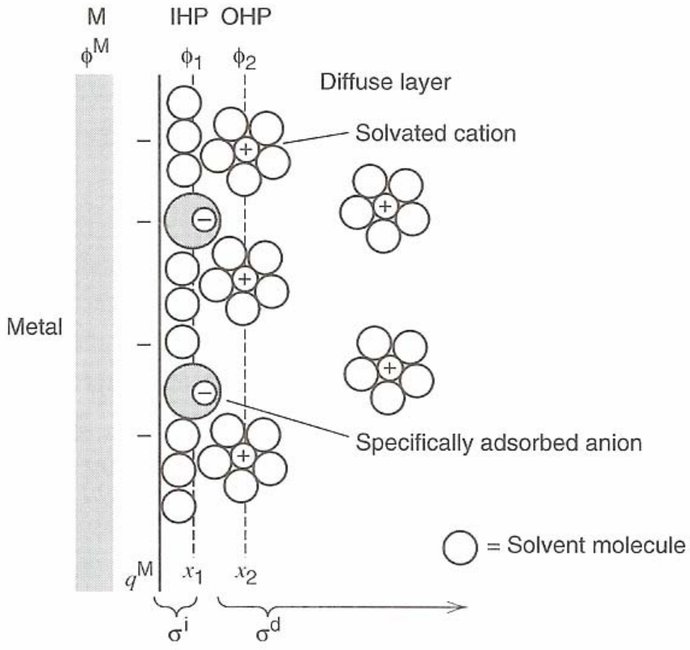


Figure 1.2.3 Proposed model of the double-layer region under conditions where anions are specifically adsorbed.

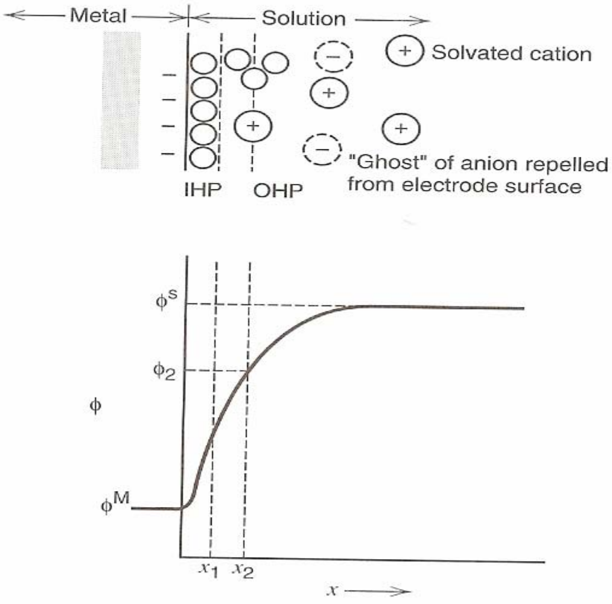
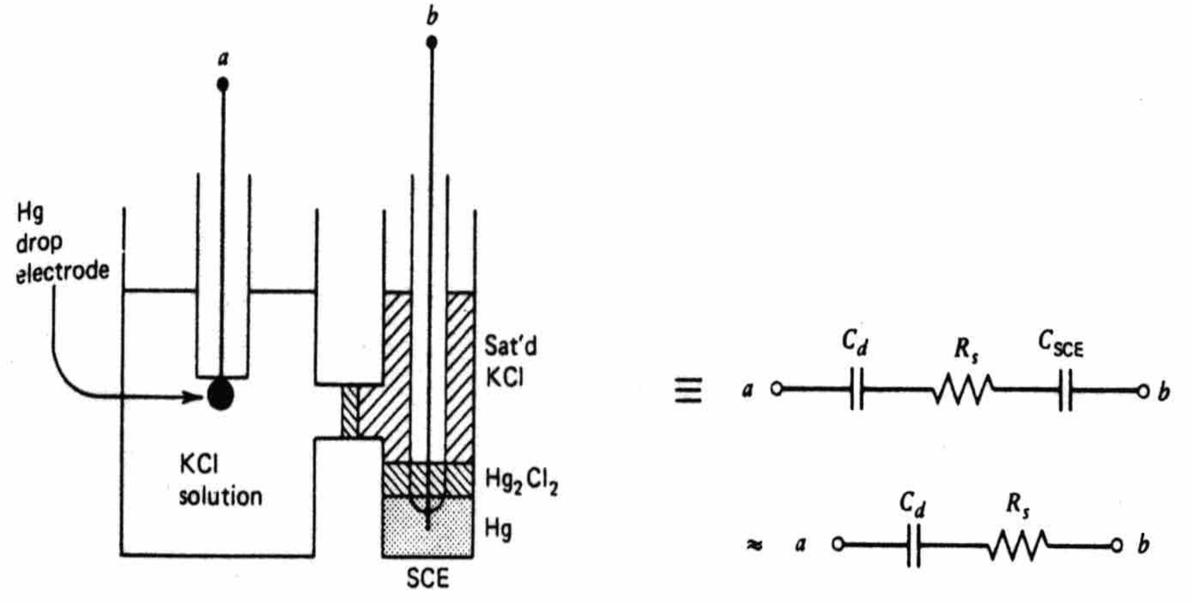


Figure 1.2.4 Potential profile across the double-layer region in the absence of specific adsorption of ions. The variable ϕ, called the *inner potential*, is discussed in detail in Section 2.2. A more quantitative representation of this profile is shown in Figure 12.3.6.

Double layer capacitance & charging current

<u>IPE and ideal reversible electrode</u>

e.g.,) Hg/K$^+$, Cl-/SCE, Hg: ideal polarized electrode



C_{SCE}, C_d: capacitances of SCE and double layer, R_s: solution resistor

$C_T = C_{SCE}C_d/(C_{SCE} + C_d)$, $C_{SCE} >> C_d \rightarrow C_T \approx C_d \rightarrow$ RC circuit

- Consider several common electrical perturbations

(a) Voltage (or potential) step:

potential step: E, E_C of capacitor, E_R of resistor

$$q = C_d E_C$$

$$E = E_R + E_C = iR_s + q/C_d$$

$$i = dq/dt$$

$$dq/dt = -q/(R_s C_d) + E/R_s$$

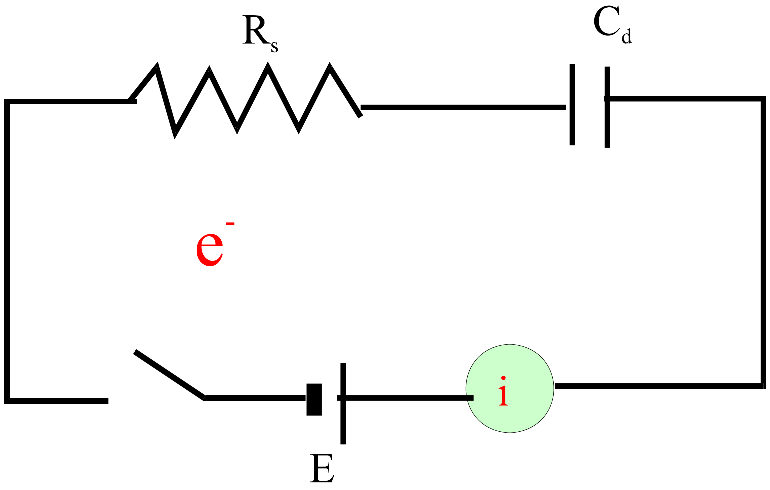
$$q = 0 \text{ at } t = 0 \rightarrow q = EC_d[1 - \exp(-t/R_s C_d)]$$

By differentiating,

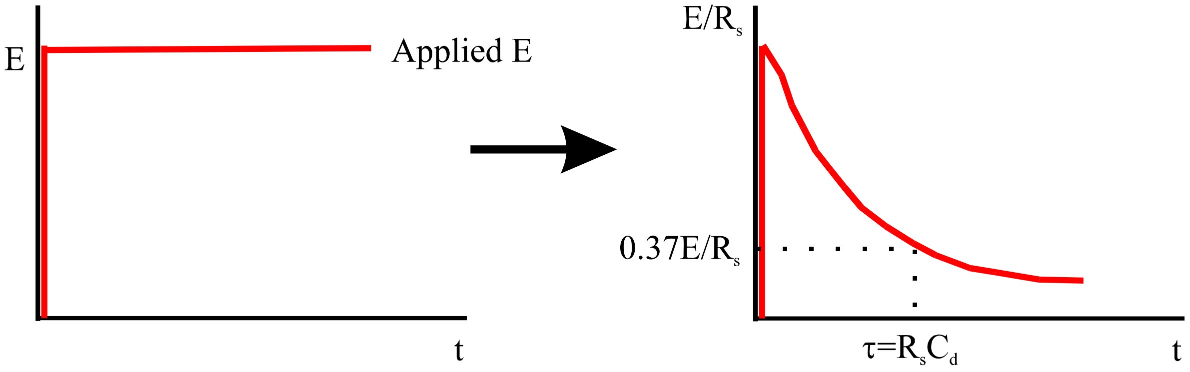
$$I = (E/R_s)\exp(-t/R_s C_d)$$

At time constant $\tau = R_s C_d$ → current for charging the double layer capacitance drops to

37 % at $\tau = t$, 5 % at $\tau = 3t$



e.g.,) $R_s = 1\ \Omega$, $C_d = 20\ \mu F$, $\tau = 20\ \mu sec$ → double layer charging is 95 % complete in 60 μsec



(b) Current step

$E = E_R + E_C = iR_s + q/C_d$

$q = \int idt$ and i is constant

$$E = iR_s + (i/C_d)\int dt$$

$$E = i(R_s + t/C_d)$$

- Potential increases linearly with time

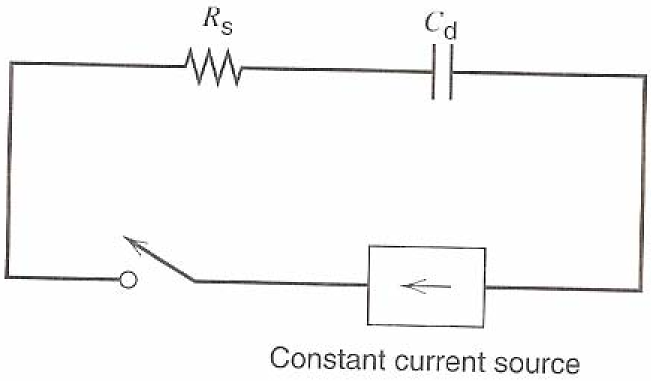


Figure 1.2.8 Current step experiment for an *RC* circuit.

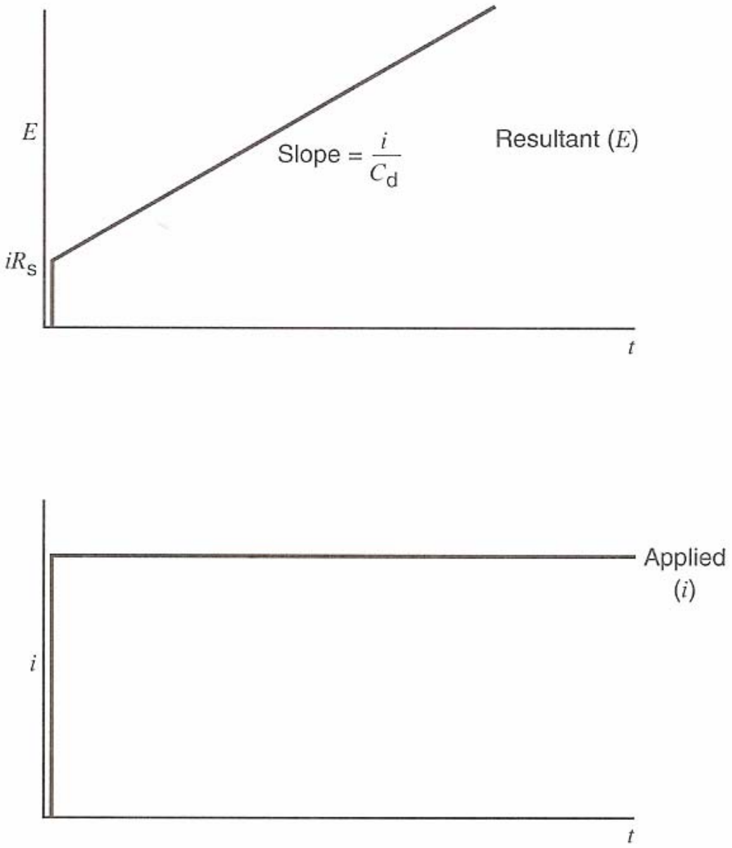


Figure 1.2.9 *E-t* behavior resulting from a current step experiment.

(c) Potential sweep

Linear potential sweep with a sweep rate v (in V/s)

$$E = vt$$

$E = E_R + E_C = iR_s + q/C_d$

$$vt = R_s(dq/dt) + q/C_d$$

If $q = 0$ at $t = 0$, $\quad i = vC_d[1 - \exp(-t/R_sC_d)]$

- Current rises from 0 and attains a steady-state value (vC_d): <u>measure C_d</u>

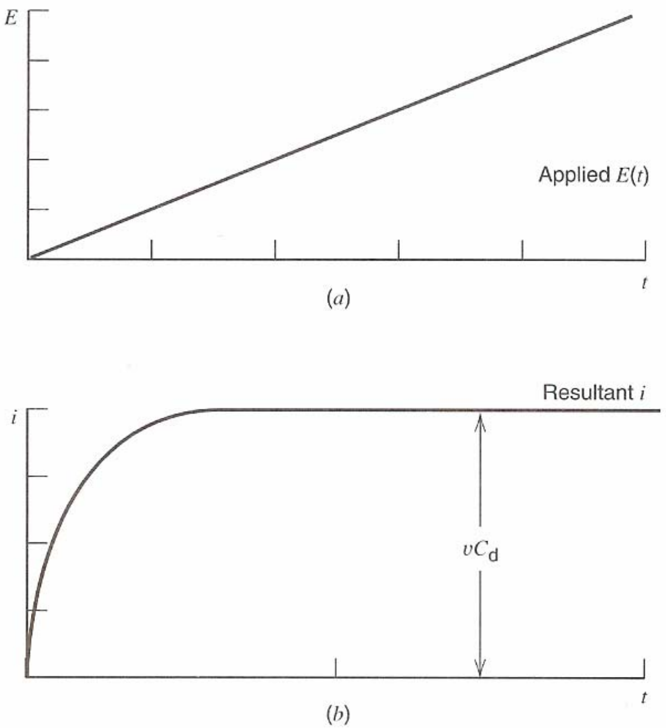


Figure 1.2.10 Current-time behavior resulting from a linear potential sweep applied to an *RC* circuit.

- Cyclic linear potential sweep

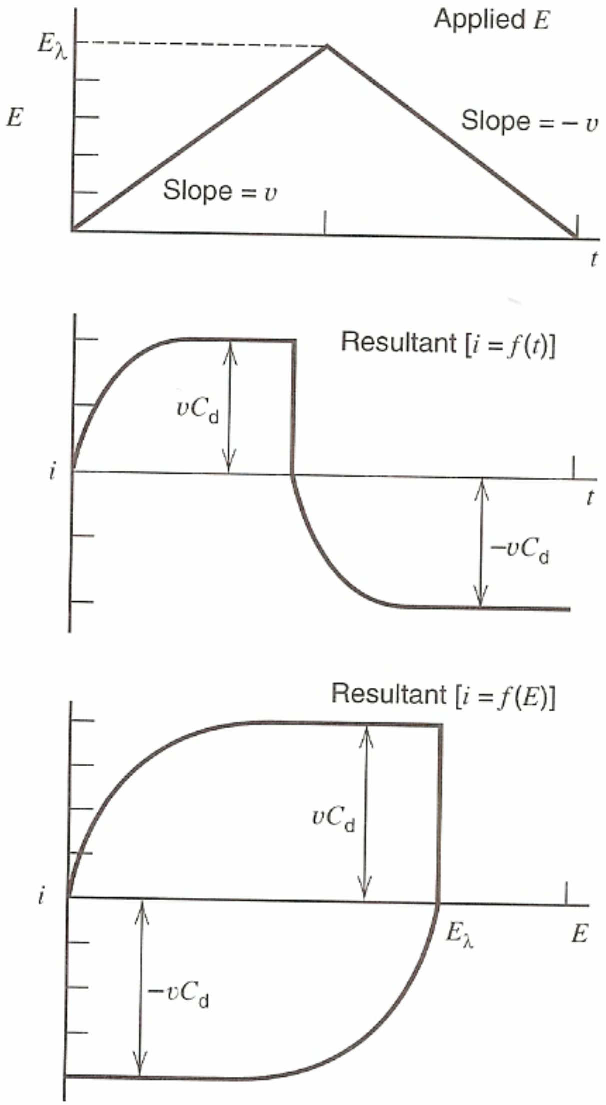


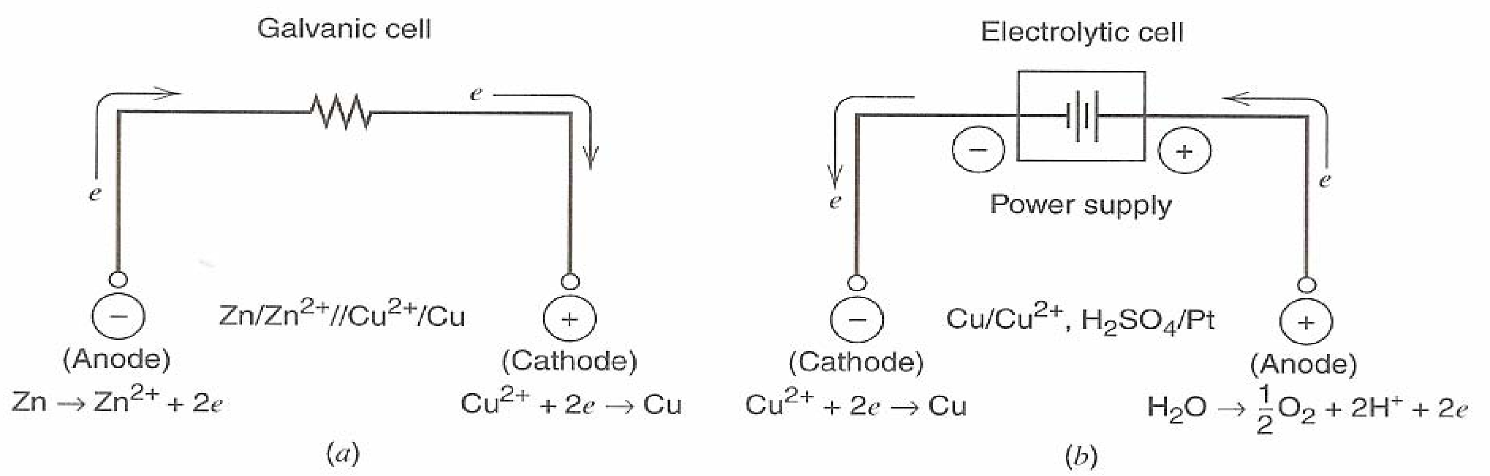
Figure 1.2.11 Current-time and current-potential plots resulting from a cyclic linear potential sweep (or triangular wave) applied to an *RC* circuit.

Faradaic processes & rates of electrode reactions

Electrochemical cells (faradaic current are flowing): types and definitions

(i) Galvanic cell: reactions occur spontaneously at the electrodes when they are connected externally by a conductor. Converting chemical energy into electrical energy. e.g., primary battery, secondary battery (discharging (방전)), fuel cell

(ii) Electrolytic cell: reactions are effected by an external voltage. Electrical energy to chemical reactions. e.g., electrolytic syntheses, electrorefining (e.g., copper), electroplating, secondary battery (charging(충전시))



(*a*)

(*b*)

Figure 1.3.1 (*a*) Galvanic and (*b*) electrolytic cells.

electrons produced: **oxidation**, "**anode**"
electrons consumed; **reduction**, "**cathode**"
In galvanic cell, anode is negative with respect to the cathode
In electrolytic cell, cathode is negative with respect to the anode

The electrochemical experiment & variables in electrochemical cells

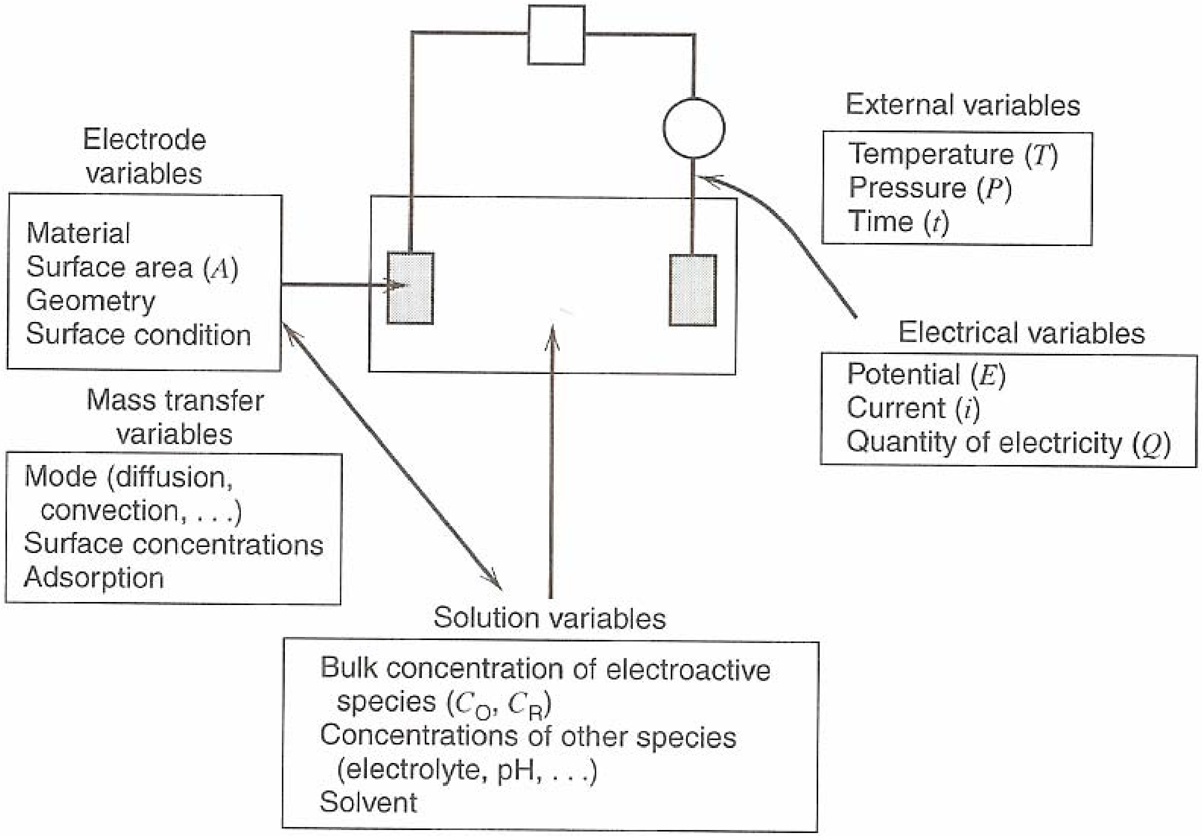


Figure 1.3.2 Variables affecting the rate of an electrode reaction.

The electrochemical experiment

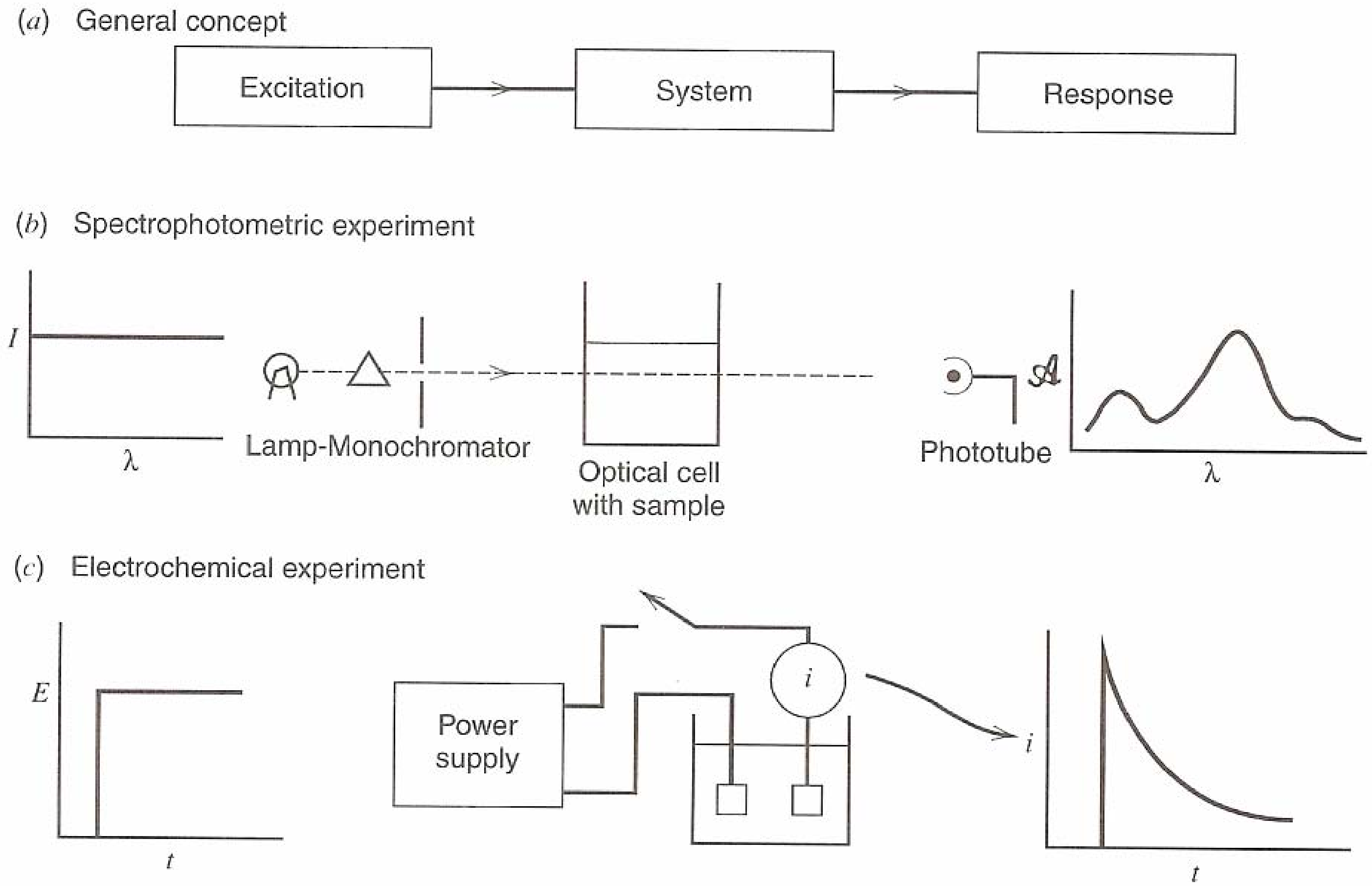


Figure 1.3.3 (*a*) General principle of studying a system by application of an excitation (or perturbation) and observation of response. (*b*) In a spectrophotometric experiment, the excitation is light of different wavelengths (λ), and the response is the absorbance ($\mathscr{A}$) curve. (*c*) In an electrochemical (potential step) experiment, the excitation is the application of a potential step, and the response is the observed i-t curve.

Current *vs.* reaction rate

$$i\ (\mathrm{A}) = dQ/dt\ (\mathrm{C/s})$$

$$Q/nF = N\ (\mathrm{mol})$$

n: # of electrons in reaction (2 for reduction of Cd^{2+})

$$\mathrm{Rate}\ (\mathrm{mol/s}) = dN/dt = i/nF$$

Electrode process: heterogeneous reaction

$$\mathrm{Rate}\ (\mathrm{mol\,s^{-1}cm^{-2}}) = i/nFA = j/nF$$

j: current density (A/cm^2)

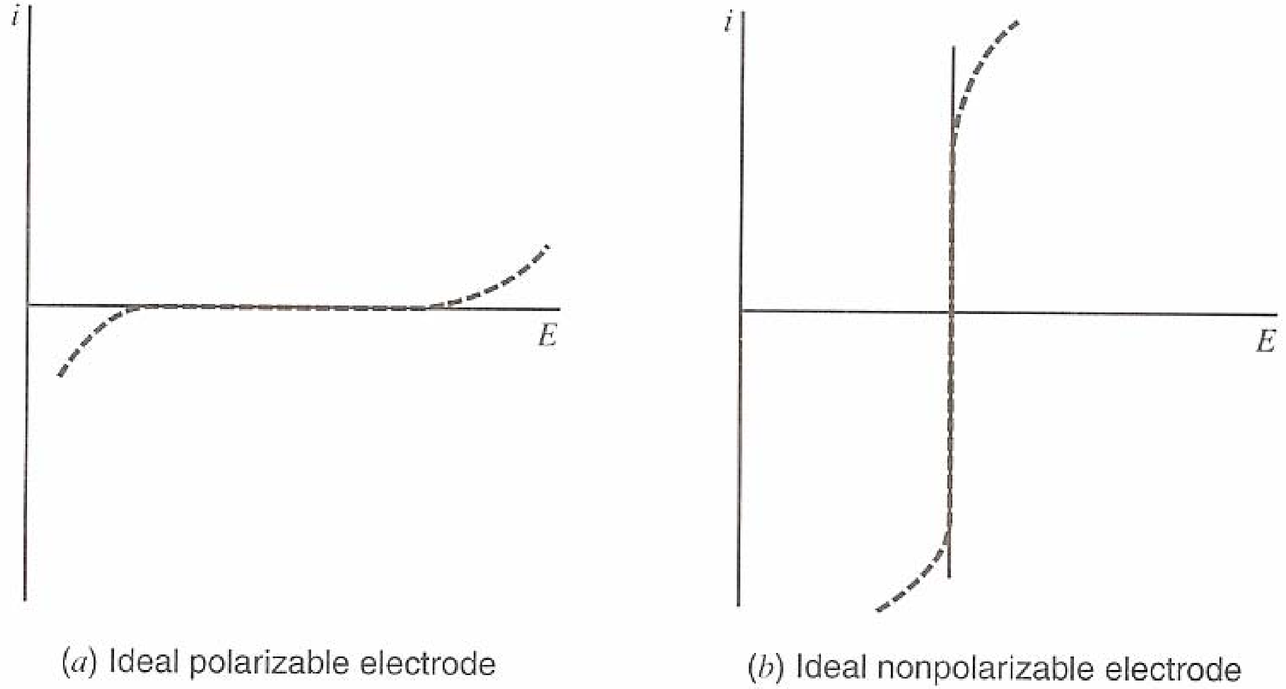
Electrode reaction: i-E curves

Polarization: departure of the cell potential from the equilibrium potential

Extent of potential measured by the overpotential: $\eta = E - E_{eq}$

Ideal polarizable electrode: a very large change in potential upon small current

Ideal nonpolarizable electrode: potential does not change upon passage of current (e.g., reference electrode)



(*a*) Ideal polarizable electrode

(*b*) Ideal nonpolarizable electrode

Figure 1.3.5 Current-potential curves for ideal (*a*) polarizable and (*b*) nonpolarizable electrodes. Dashed lines show behavior of actual electrodes that approach the ideal behavior over limited ranges of current or potential.

Factors affecting electrode reaction rate and current

1. Mass transfer
2. Electron transfer at the electrode surface
3. Chemical reactions
4. Other surface reactions: adsorption, desorption, electrodeposition

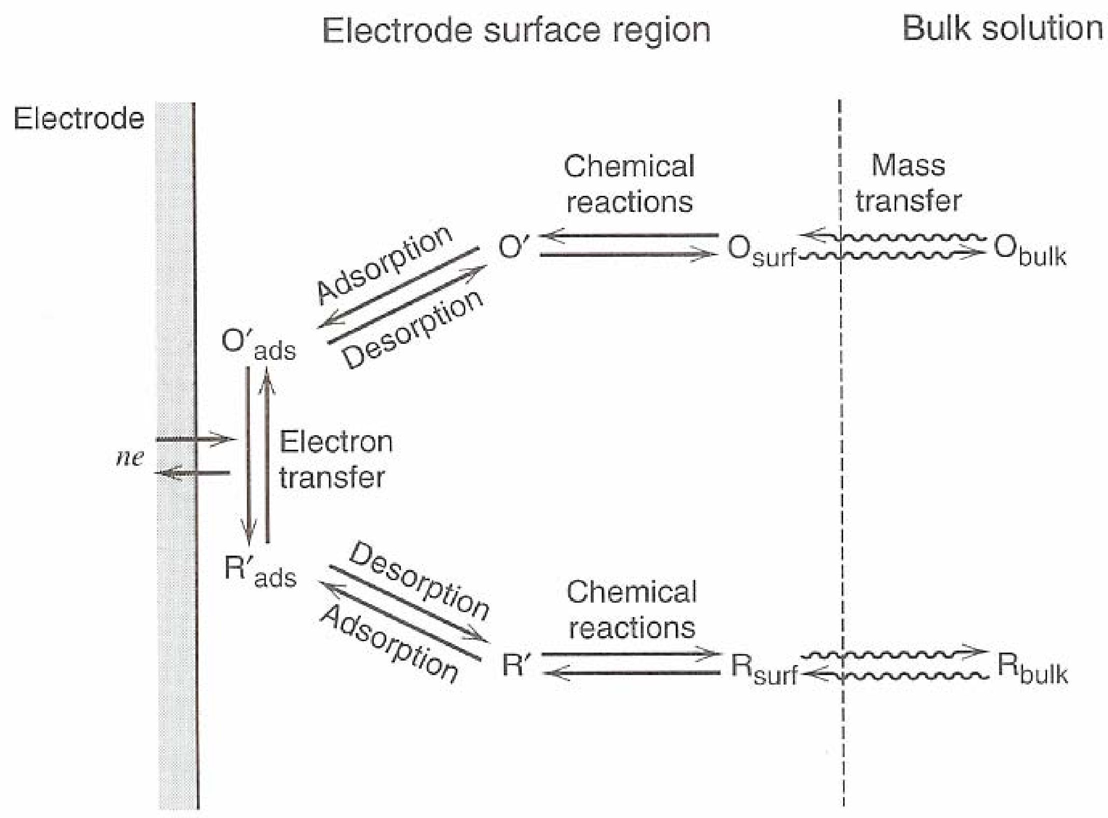


Figure 1.3.6 Pathway of a general electrode reaction.

Electrochemical cells & cell resistances

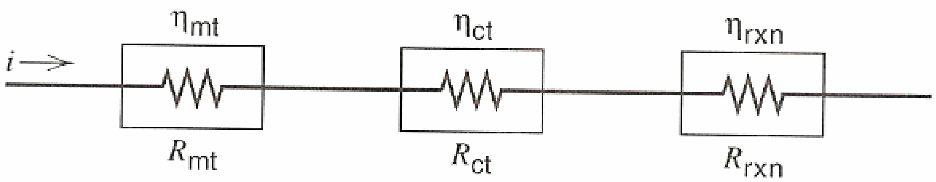


Figure 1.3.7 Processes in an electrode reaction represented as resistances.

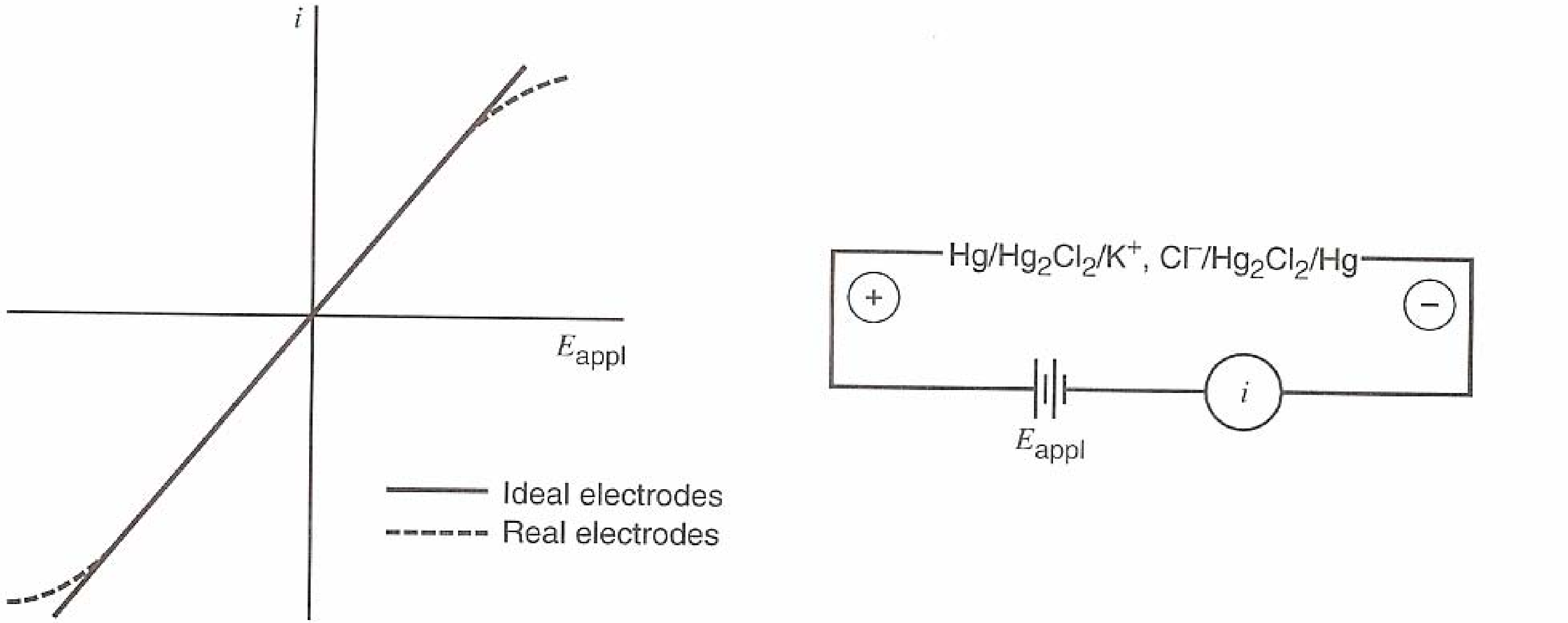
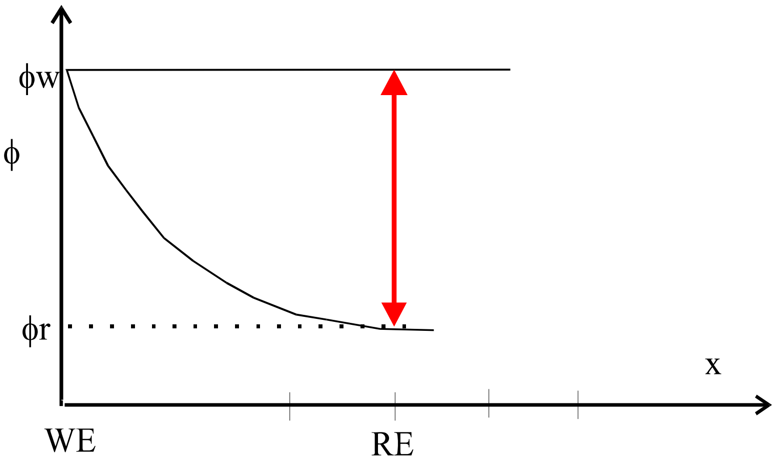
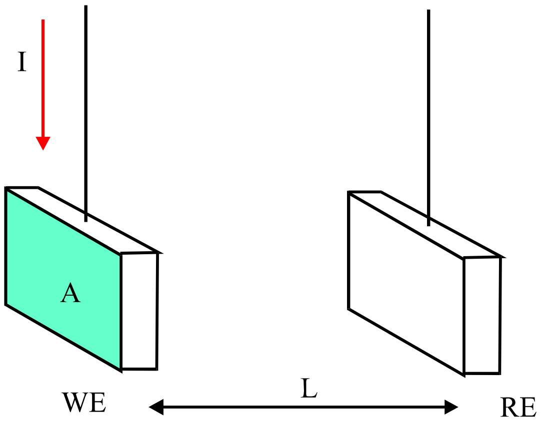


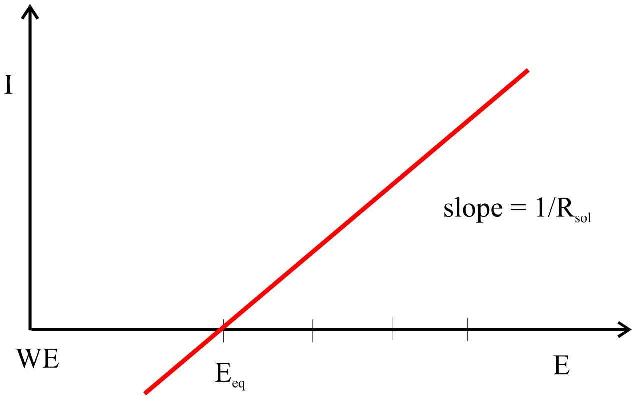
Figure 1.3.8 Current-potential curve for a cell composed of two electrodes approaching ideal nonpolarizability.

$\eta_{ohm} = IR_{sol}$, “IR drop”



$R_{sol} = L/\kappa A$





If free of charge transfer and mass transfer overpotentials, slope = $1/R_{sol}$

Electrochemistry needs to minimize η_{ohm}
κ (conductivity) ↑→ η_{ohm} ↓ (by adding extra electrolyte: "supporting electrolyte")
Closer between WE and RE
three-electrode system

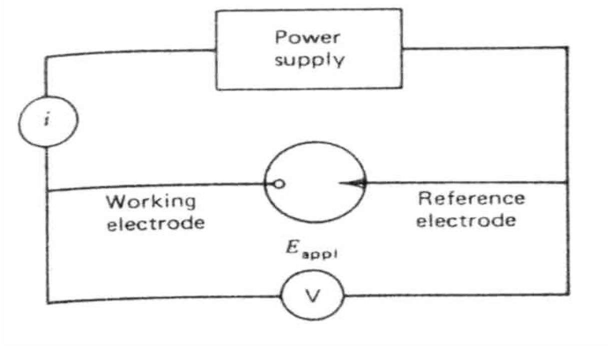
<u>two-electrode cell vs. three-electrode cell</u>

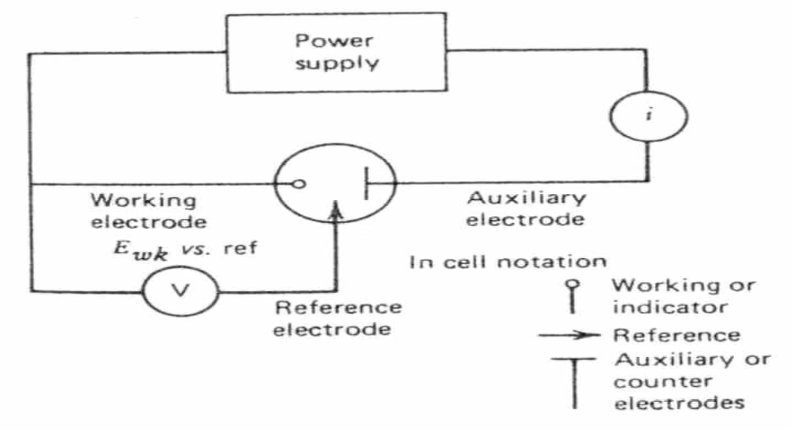
$$E_{appl} = E + iR_s = E_{eq} + \eta + iR_s$$

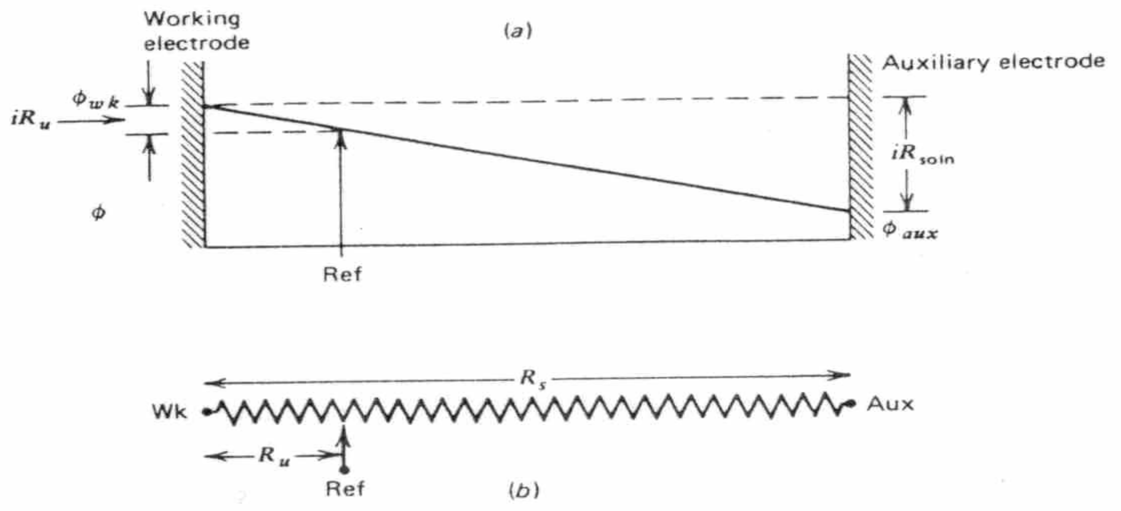
IR_s: ohmic drop in the solution (ohmic polarization) → should be minimized → short distance between working and reference electrode & three-electrode cell

Two-electrode cell: iR_s problem due to high current flow
Three-electrode cell: current between WE and auxiliary electrode(or counter electrode)
Potential measurement between WE and RE → almost no current to reference electrode

→ **Potentiostat, etc electrochemical system: three electrode system**


Power supply
i
Working electrode
Reference electrode
E_{appl}
V


Power supply
i
Working electrode
Auxiliary electrode
E_{wk} vs. ref
V
Reference electrode
In cell notation
Working or indicator
Reference
Auxiliary or counter electrodes


Working electrode
(a)
Auxiliary electrode
ϕ_{wk}
iR_u
ϕ
iR_{soln}
ϕ_{aux}
Ref
R_s
Wk
Aux
R_u
Ref
(b)

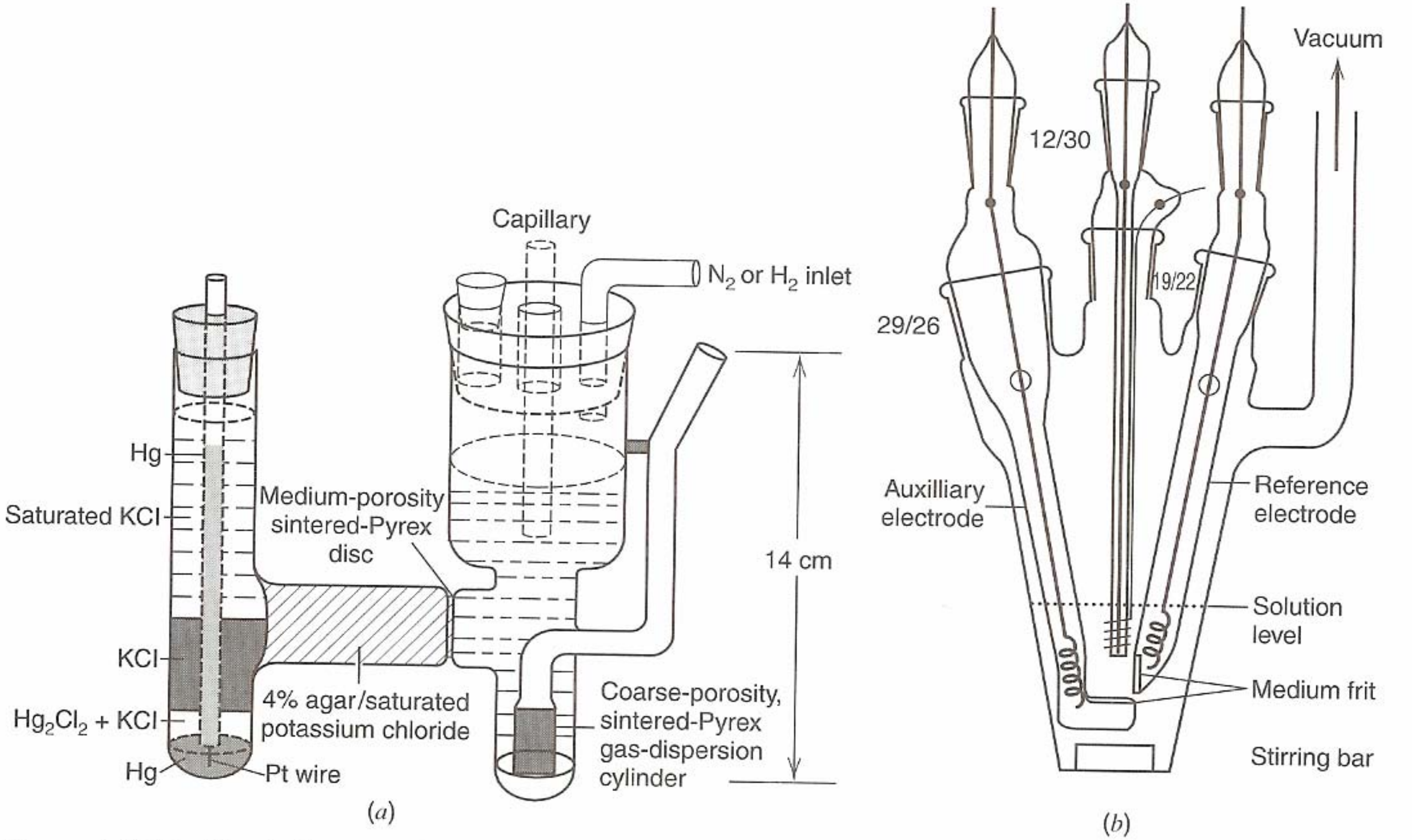


Figure 1.3.11 Typical two- and three-electrode cells used in electrochemical experiments. (*a*) Two-electrode cell for polarography. The working electrode is a dropping mercury electrode (capillary) and the N_2 inlet tube is for deaeration of the solution. [From L. Meites, *Polarographic Techniques*, 2nd ed., Wiley-Interscience, New York, 1965, with permission.] (*b*) Three-electrode cell designed for studies with nonaqueous solutions at a platinum-disk working electrode, with provision for attachment to a vacuum line. [Reprinted with permission from A. Demortier and A. J. Bard, *J. Am. Chem. Soc.*, **95**, 3495 (1973). Copyright 1973, American Chemical Society.] Three-electrode cells for bulk electrolysis are shown in Figure 11.2.2.

Mass transport-controlled reactions

Modes of mass transfer

Electrochemical reaction at electrode/solution interface: molecules in bulk solution must be transported to the electrode surface → "mass transfer"

Mass transfer-controlled reaction

$$v_{rxn} = v_{mt} = i/nFA$$

Modes for mass transport:

(a) Migration: movement of a charged body under the influence of an electric field (a gradient of electric potential)

(b) Diffusion: movement of species under the influence of gradient of chemical potential (i.e., a concentration gradient)

(c) Convection: stirring or hydrodynamic transport

Nernst-Planck equation (diffusion + migration + convection)

$$J_i(x) = -D_i(\partial C_i(x)/\partial x) - (z_iF/RT)D_iC_i(\partial \phi(x)/\partial x) + C_iv(x)$$

Where $J_i(x)$; the flux of species i ($\mathrm{mol sec^{-1} cm^{-2}}$) at distance x from the surface, D_i; the diffusion coefficient ($\mathrm{cm^2/sec}$), $\partial C_i(x)/\partial x$; the concentration gradient at distance x, $\partial \phi(x)/\partial x$; the potential gradient, z_i and C_i; the charge and concentration of species i, v(x); the velocity (cm/sec)

Steady state mass transfer

steady state, $(\partial C/\partial t) = 0$; the rate of transport of electroactive species is equal to the rate of their reaction on the electrode surface

In the absence of migration (excess supporting electrolyte),

$$O + ne^- = R$$

The rate of mass transfer,

$$v_{mt} \propto (\partial C_O(x)/\partial x)_{x=0} = D_O(C_O{}^b - C_O{}^s)/\delta$$

where x is distance from the electrode surface & δ: diffusion layer

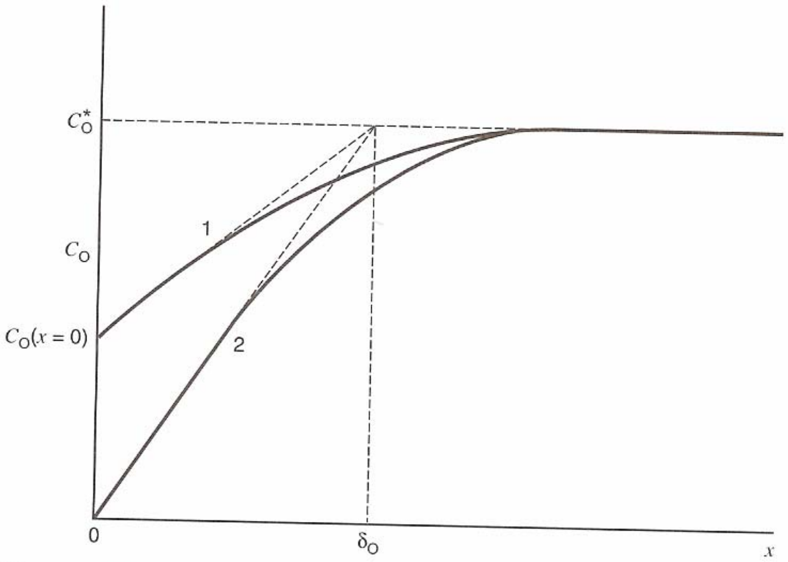


Figure 1.4.1 Concentration profiles (solid lines) and diffusion layer approximation (dashed lines). $x = 0$ corresponds to the electrode surface and δ_O is the diffusion layer thickness. Concentration profiles are shown at two different electrode potentials: (*1*) where $C_O(x = 0)$ is about $C_O^*/2$, (*2*) where $C_O(x = 0) \approx 0$ and $i = i_l$.

$$v_{mt} = m_O[C_O^b - C_O^s]$$

where C_O^b is the concentration of O in the bulk solution, C_O^s is the concentration at the electrod surface

m_O is “mass transfer coefficient (cm/s)” ($m_O = D_O/ \delta$)

$$i = nFAm_O[C_O^b - C_O^s]$$

$$i = -nFAm_R[C_R^b - C_R^s]$$

largest rate of mass transfer of O when $C_O^s = 0 \rightarrow$ "limiting current"

$$\mathbf{i_{l,c} = nFAm_O C_O^b}$$

Maximum rate when limiting current flows

$$C_O^s/C_O^b = 1 - (i/i_{l,c})$$

$$C_O^s = [1 - (i/i_{l,c})]\,[\, i_{l,c}/nFAm_O] = (i_{l,c} - i)/(nFAm_O)$$

C_O^s varies from C_O^b at $i = 0$ to negligible value at $i = i_l$

If kinetics of electron transfer are rapid, the concentrations of O and R at the electrode surface are at equilibrium with the electrode potential, as governed by the Nernst equation for the half-reaction

$$\mathbf{E = E^{0\prime} + (RT/nF)\ln(C_O^s/C_R^s)}$$

$E^{0\prime}$: formal potential (activity coeff.), cf. E^0 (standard potential)

(a) R initially absent

When $C_R^b = 0$, $\quad C_R^s = i/nFAm_R$

$$C_O^s = (i_{l,c} - i)/(nFAm_O)$$

$$\mathbf{E = E^{0\prime} - (RT/nF)\ln(m_O/m_R) + (RT/nF)\ln(i_{l,c} - i/i)}$$

i-E plot

When $i = i_{l,c}/2$, $E = E_{1/2} = E^{0\prime} - (RT/nF)\ln(m_O/m_R)$

$E_{1/2}$ is <u>independent of concentration & characteristic of O/R system</u>

$$E = E_{1/2} + (RT/nF)\ln(i_{l,c} - i/i)$$

Plot of E vs. $\log(i_{l,c} - i/i)$: straight, slope = 2.3RT/nF (or 59.1/n mV^{-1} at 25ºC)

E intercept = $E_{1/2}$ (when $m_O \sim m_R \rightarrow E_{1/2} \sim E^{0\prime}$)

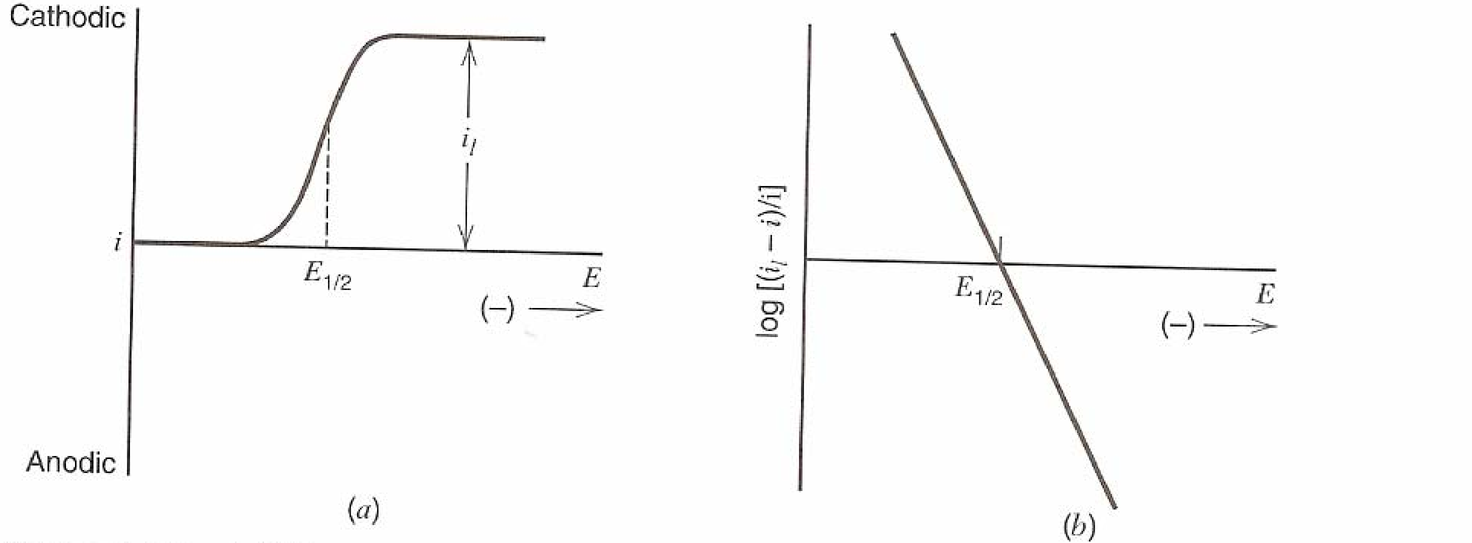


Figure 1.4.2 (*a*) Current-potential curve for a nernstian reaction involving two soluble species with only oxidant present initially. (*b*) $\log[(i_l - i)/i]$ *vs*. *E* for this system.

(b) Both O and R initially present

Same method,

$$C_R^s/C_R^b = 1 - (i/i_{l,a})$$
$$i_{l,a} = -nFAm_RC_R^b$$
$$C_R^s = -[1 - (i/i_{l,a})]\,[\,i_{l,a}/nFAm_R] = -(i_{l,a} - i)/(nFAm_R)$$

Put these equations to $E = E^{0\prime} + (RT/nF)\ln(C_O^s/C_R^s)$

$$\mathbf{E = E^{0\prime} - (RT/nF)\ln(m_O/m_R) + (RT/nF)\ln[(i_{l,c} - i)/(i - i_{l,a})]}$$

When $i = 0$, $E = E_{eq}$ and the system is at equilibrium

Deviation from E_{eq}: concentration overpotential

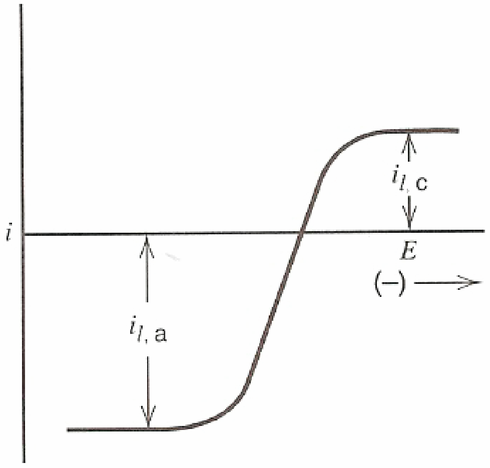


Figure 1.4.3 Current-potential curve for a nernstian system involving two soluble species with both forms initially present.

(c) R insoluble

Suppose R is a metal (plating),

$a_R = 1$, $E = E^{0\prime} + (RT/nF)\ln C_O^s$

Using $C_O^s/C_O^b = 1 - (i/i_{l,c})$

$$\mathbf{E = E^{0\prime} + (RT/nF)\ln C_0^b + (RT/nF)\ln[(i_l - i)/i_l]}$$

When $i = 0$, $E = E_{eq} = E^{0\prime} + (RT/nF)\ln C_O^b$

Deviation from E_{eq}: concentration overpotential

$$\eta_{conc} = E - E_{eq} = (RT/nF)\ln[(i_l - i)/i_l]$$

When $I = i_l$, $\eta_{conc} \to \infty$ (complete concentration polarization)

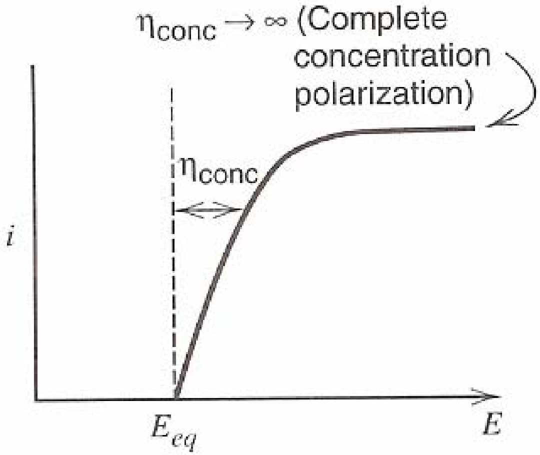


Figure 1.4.4 Current-potential curve for a nernstian system where the reduced form is insoluble.

$$1 - i/i_l = \exp(nF\,\eta_{conc}/RT)$$

$e^x = 1 + x + \ldots = 1 + x$ (when x is small)

At small deviation of potentials from E_{eq}, i- η_{conc} is linear

$$\eta_{conc} = -RTi/nFi_l$$

Mass transfer resistance, $R_{mt} = RT/nF|i_l|$

Non-steady state mass transport: diffusion control

Time-dependent (transient) phenomena

The rate of diffusion depends on the concentration gradients

$$J = -D(\partial C/\partial x) \quad \text{Fick's first law}$$

D: diffusion coefficient (cm^2/sec)

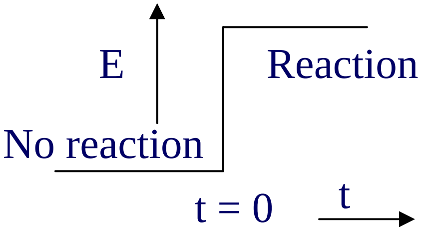
The variation of concentration with time due to diffusion → Fick's second law

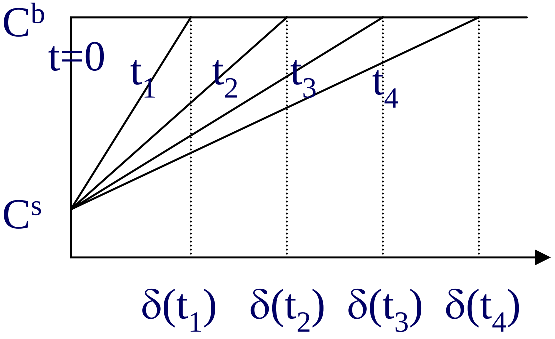
$$\partial C/\partial t = -D(\partial^2 C/\partial x^2) \qquad \text{1-D}$$

$$J = -D(\partial C/\partial x) = i/nFA$$

$$v_{mt} = D(\partial C/\partial x) = D_O(C_O^b - C_O^s)/\delta(t) = i/nFA$$

Time-dependent, applying potential step E





Moles of species in diffusion layer = $\int i dt/nF \approx [C^b - C^s](A\delta(t)/2)$

Differentiating,

$i/nF = [C^b - C^s](A d\delta(t)/2dt) = DA(C^b - C^s)/\delta(t)$

$d\delta(t)/dt = 2D/\delta(t)$, $\delta(t) = 0$ at $t = 0$

$\delta(t) = 2\sqrt{(Dt)}$

$$i/nFA = (D^{1/2}/2t^{1/2})\,[C^b - C^s]$$

diffusion layer grows with $t^{1/2}$ and current decays with $t^{-1/2}$

In the absence of convection: current continues to decay
In a convection system: approaches to steady-state value

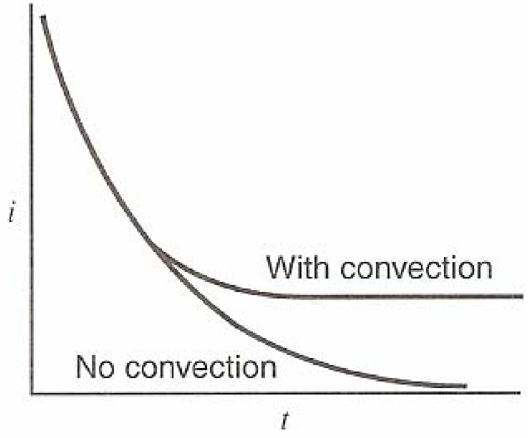


Figure 1.4.6 Current-time transient for a potential step to a stationary electrode (no convection) and to an electrode in stirred solution (with convection) where a steady-state current is attained.