Biosensors

B. Tech.

Course No.: EEL 3050

L-T-P [C]: 3–0–2 [4]

Prof. AJAY AGARWAL

ELECTRICAL ENGINEERING

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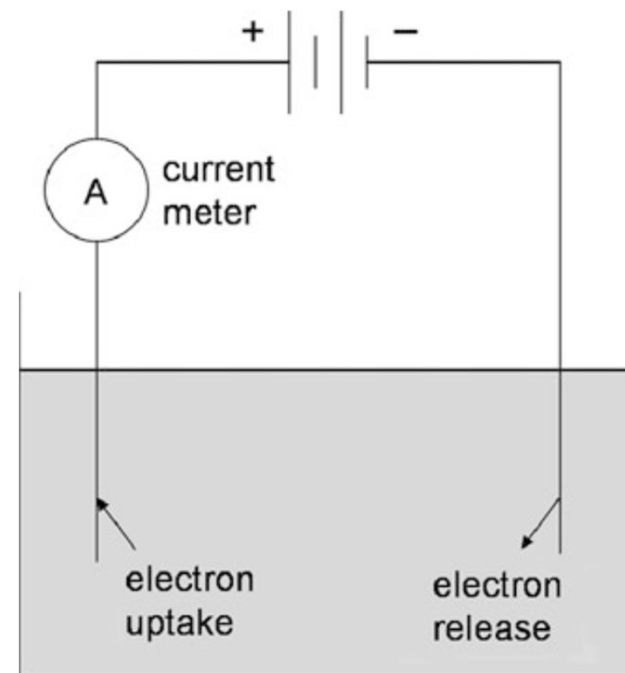
Lecture 22 dated 3rd Oct. 2024

Electrochemical Sensors:

- The top two most popular **transducers** for biosensors are **optical & electrochemical,** followed by piezoelectric & thermal.
- **Electrochemical sensors** are electrochemical transducers, where the **conc. of ions or chemicals** are converted into
  - electrical **voltage** (potentiometric),
  - electrical **current** (amperometric), or
  - electrical **resistance/ conductance** (conductometric).
- If electrochemical sensors are used together with bioreceptors (**enzymes** or **antibodies**), they become electrochemical biosensors.

Electrolytic and Electrochemical Cells:
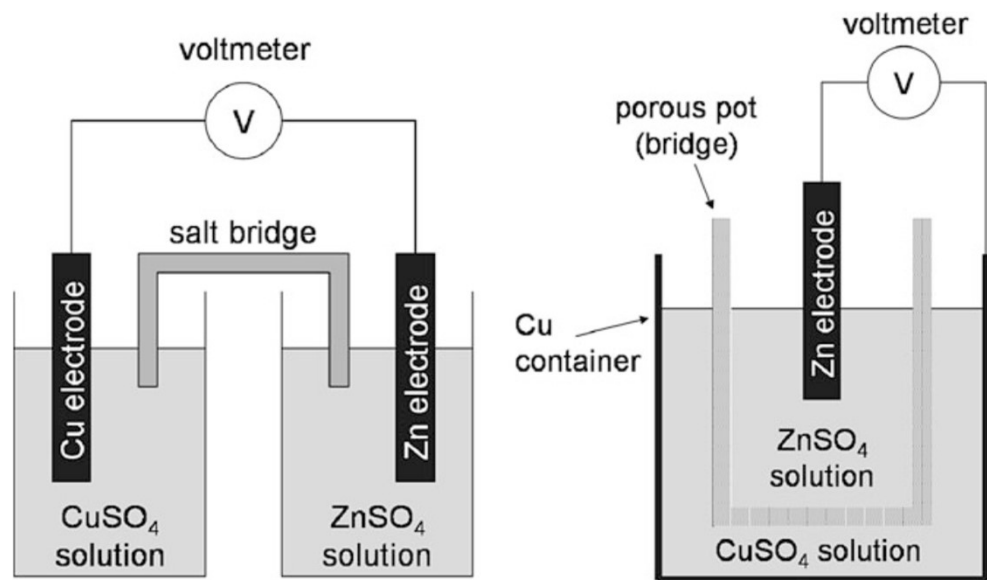
- An electrolytic cell decomposes **ionic chemical compounds** by applying **voltage** to its solution.
- a typical electrolytic cell has two electrodes (metal rods) inserted into a solution of metal salts (electrolytes).
- Electrons are taken from metal ions at one electrode (oxidation) and are released to metal ions at the other electrode (reduction).
- Whole reaction is called redox (reduction + oxidation) reaction.
- The current that flows between the electrodes depends not only on the voltage that is applied, but also on the electrical properties of the solution.
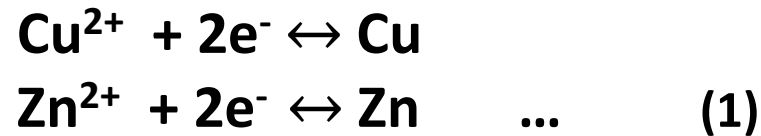

An electrolytic cell

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Electrochemical Cells:

- An electrochemical cell is a **device** that **generates** electrical voltage and current.
- Two electrodes are inserted into electrolytes that are separated by a **salt bridge** (**Galvanic** cell), or a **semi-permeable membrane** that replaces the function of a salt bridge (Daniell cell)
- The cells are identical to each other
- Each side of the cells is referred to as a half-cell.
- The cells are described as:

Cu (s) | CuSO4 (aq) | | ZnSO4 (aq) | Zn (s)

Electrochemical cells: galvanic (left) and Daniell (right) cells

The following chemical formulas describe the equilibrium of both metal ions:

$$Cu^{2+} + 2e^- \leftrightarrow Cu$$

$$Zn^{2+} + 2e^- \leftrightarrow Zn \qquad \ldots \qquad (1)$$

Combining these two formulas yields:

$$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+} \qquad \ldots \qquad (2)$$

Here, forward reaction is spontaneous i.e. the change in Gibbs free energy is negative or $\Delta G < 0$. ΔG is defined as

$$\Delta G = -n.F.\Delta E \qquad \ldots \qquad (3)$$

where,
n number of electrons
F Faraday constant = 96,487 C mol^{-1}
ΔE electrical potential difference (V)

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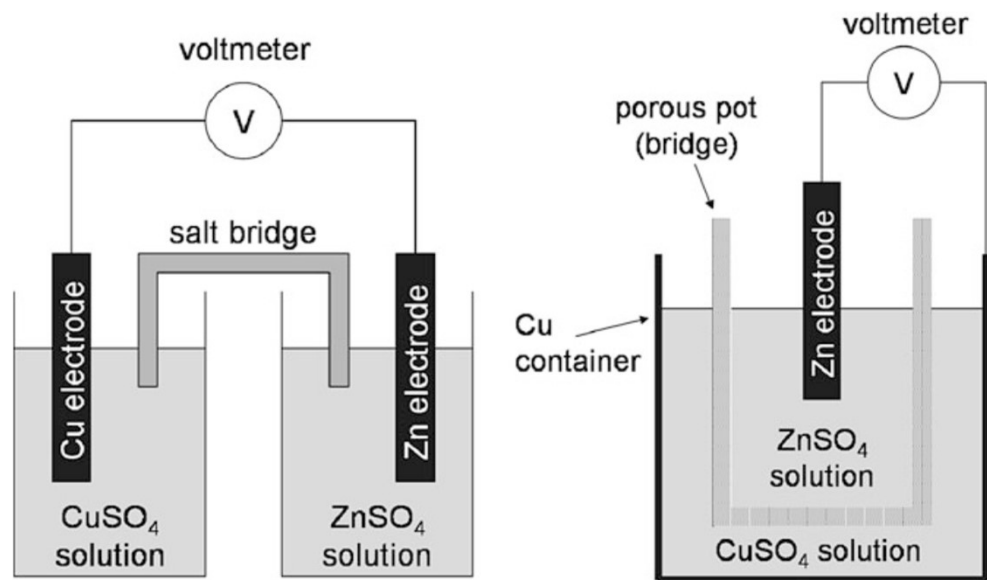
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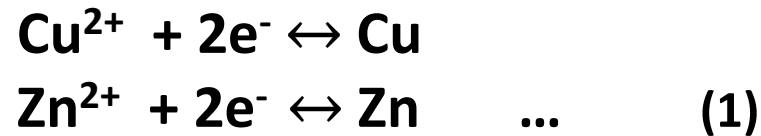
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Lecture 23 dated 16th Oct. 2024

Electrochemical Cells:

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$$\Delta G = -n.F.\Delta E \qquad \ldots \qquad (3)$$

where,
n number of electrons
F Faraday constant = 96,487 C mol^{-1}
ΔE electrical potential difference (V)

We also know that

$$\Delta G = RT \ln K \qquad \ldots \qquad (4)$$

where,
R gas constant = $8.3145 \text{ J K}^{-1} \text{ mol}^{-1}$
T temperature (K)
K equilibrium constant

Combining above Eqs. 3 & 4 gives

$$\Delta E = -RT/ nF \times \ln K \qquad \ldots \qquad (5)$$

Applying Eq. 5 to the left-side half-cell of the above electrochemical cell: Cu2+ + 2e– = Cu.

For this reaction, n is 2, and K is defined by the **activity ratio**

$$K = a_{Cu} / a_{Cu2+} \qquad \dots \qquad (6)$$

The activity of a **pure liquid or pure solid** is 1 (a_{Cu} = 1), and the activity of **metal ions can be approximated** by their **molar concentration** ($a_{Cu2+} = [Cu^{2+}]$).

ΔE can also be replaced with $E - E°$, where **E° refers to the standard electrode potential** (reference or ground potential).

Plugging all of these into Eq. 10.5 yields

$$E = E^\circ - \frac{RT}{2F} \times \ln \frac{1}{[Cu^{2+}]}$$

or

$$E = E^\circ + \frac{RT}{2F} \times \ln [Cu^{2+}] \quad \ldots \quad (7)$$

Or more generally,

$$E = E^\circ + \frac{RT}{nF} \times \ln [M^+] \quad \ldots \quad (8)$$

which is called the Nernst equation. Notice that the **potential** (voltage) is **linearly proportional** to the **logarithm of ionic concentration**.

We can also convert the natural log into a common log using ln X = 2.303 log X. By plugging in universal constants for R, F, and a room temperature of 25 °C (=273.15 K) into Eq. 8

$$E = E° + 0{:}059/ n \; \times \log [M^+] \qquad ... \qquad (9)$$

For the entire cell,

$$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+} \qquad ... \qquad (2)$$

$$K = a_{Cu}\, a_{Zn^{2+}}/ a_{Cu^{2+}}\, a_{Zn} = a_{Zn^{2+}}/ a_{Cu^{2+}} \approx [Zn^{2+}]/ [Cu^{2+}] \quad ... \qquad (10)$$

as the activities of pure solids are 1 ($a_{Cu} = a_{Zn} = 1$).

Plugging Eq. 10 into Eq. 5 and repeating the above procedure yields

$$E = E^\circ + 0{:}059/\,n \ \times \log\,[M^+] \qquad \dots \qquad (9)$$

For the entire cell,

$$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+} \qquad \dots \qquad (2)$$

$$K = a_{Cu}\, a_{Zn^{2+}}/\,a_{Cu^{2+}}\, a_{Zn} = \ a_{Zn^{2+}}/\,a_{Cu^{2+}} \approx [Zn^{2+}]/\,[Cu^{2+}] \ \dots \qquad (10)$$

as the activities of pure solids are 1 ($a_{Cu} = a_{Zn} = 1$). Plugging Eq. 10 into Eq. 5 & repeating the above procedure yields

$$E = E^\circ + 0.059/\,n \ \times \log[Zn^{2+}]/\,[Cu^{2+}] \qquad \dots \qquad (11)$$

For a Galvanic or Daniell cell, the standard electrode potential for the entire reaction can be calculated from those of half-cells. The forward reaction of Eq. 2 is spontaneous, which can be split into

$$Cu^{2+} + 2e^- \rightarrow Cu \qquad E° = + 0.34V$$
$$+)\ Zn \rightarrow Zn^{2+} + 2e^- \qquad E° = -(-0.762)V = 0.762V$$
_____
$$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+} \qquad E° = +0.34 + 0.762 = +1.102V \qquad ... \qquad (12)$$

Half-cell reaction	$E°$ (V)
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	–0.762
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	–0.44
$2H^+ + 2e^- \rightleftharpoons H_2$	0.0
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34
$Hg_2Cl_2 + 2e^- \rightleftharpoons 2Hg + 2Cl^-$	+0.242

Select standard electrode potentials ($E°$)

If, $[Cu^{2+}] = [Zn^{2+}] = 0.1M$,

$E = + 1.102 + (0.059/2) \times \log 0.1/ 0.1 = +1.102V$

If, $[Cu^{2+}] = 1M$ & $[Zn^{2+}] = 0.1M$,

$E = + 1.102 + (0.059/2) \times \log 1/ 0.1 = +1.13V$

Typical Galvanic or Daniell cells generate ~1.1V unless the ion **concentrations** of **copper & zinc** are significantly different by a **few orders of magnitude.**

Questions and Discussion?

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Lecture 24 dated 17th Oct. 2024

Amperometric Biosensors:

- Enzymes are popular bioreceptors in identifying & quantifying their counterpart, substrates.
- Enzymes are typically proteins, while substrates are generally small chemicals.
- A good example is the use of glucose oxidase (GOx; the names of **enzymes** usually **end** with **-ase**) to quantifying glucose.
- GOx converts glucose into gluconic acid while generating hydrogen peroxide as a byproduct.
- During this reaction, GOx itself is reduced.

$$\text{Glucose} + O_2 \xrightarrow{\text{GOx}} \text{gluconic acid} + H_2O_2$$
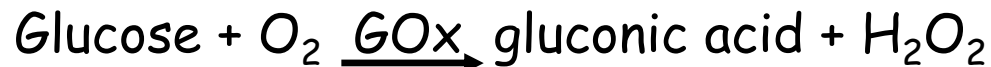
- An appropriate electron mediator can be added, to regenerate GOx back into its oxidized (i.e., active) form

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- This **redox cycle generates** electrical **current** (with no voltage applied; not practical) or **change in electrical current** (with voltage applied).
- Since the **quantification** is made through measuring current, it is referred to as amperometric electrochemical biosensor or simply amperometric biosensor.
- This type of biosensor is particularly popular in measuring blood glucose level from diabetic patients.

Conductometric Biosensors:

- In conductometric biosensors, conductivity (inverse of resistivity) is measured to monitor the redox reactions of enzymatic oxidation.
- It is similar to amperometric biosensors; the only difference is the current measured in amperometry & conductivity in conductometry.

$$resistivity = RA/ l$$

where,
R resistance (Ω)
A cross-sectional area of a specimen (cm^2)
l length of a specimen (cm)
The typical unit of resistivity is $\Omega \cdot cm$.

- Conductivity is defined as: $conductivity = l / RA$

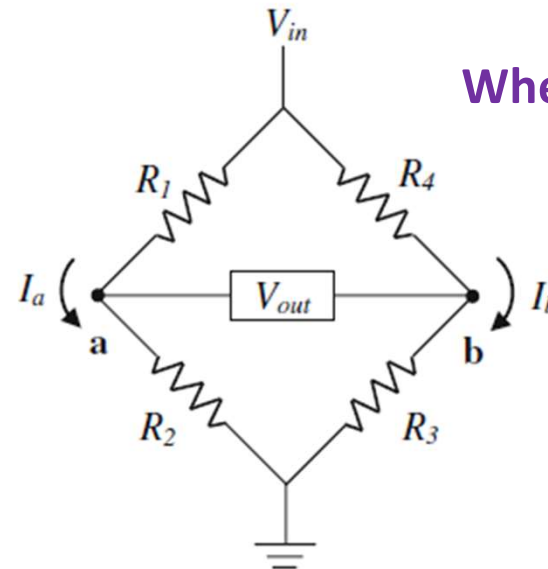The unit of conductivity is $\Omega^{-1} \cdot cm^{-1}$, or $S \cdot cm^{-1}$ (the standard unit),
S, is called Siemens.

Conductivity cell:
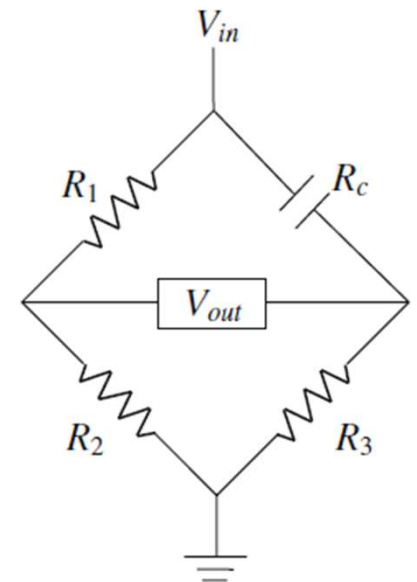
- Both resistivity & conductivity are measured with a conductivity cell,
- It has **two metal plates separated** by a **certain distance**.
- Here, two 1 cm^2 (= A) metal plates separated by 1 cm (= l).
- The ratio l/A is referred to as cell constant, & it is 1 cm^{-1} (= 1 cm/ 1 cm^2)
- This **conductivity cell** is used to **quantify** the extent of the enzymatic redox reactions.
- Such conductivity changes, however, are generally too small to be detected.

Two metal plates
e.g. 1 cm^2 each
1 cm apart

direct contact
w/ liquid
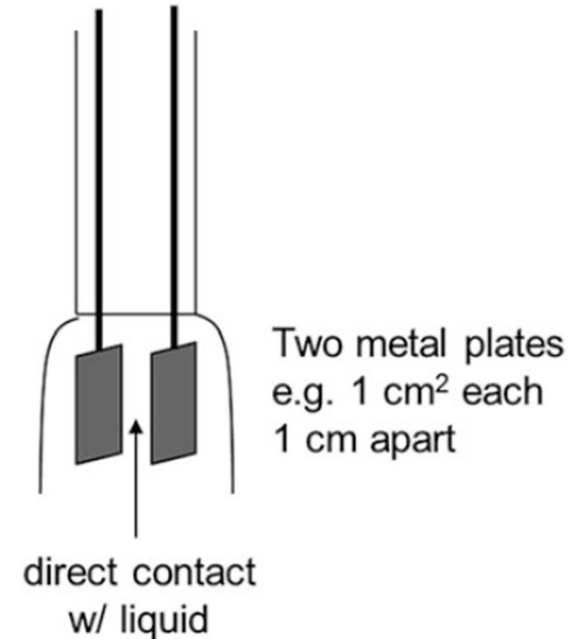
- Conductivity is inverse of resistivity
- the small resistance change can be measured by a Wheatstone bridge
- If a conductivity cell replace the unknown resistor (R4) from a Wheatstone bridge, it is called conductivity bridge,
- a basic platform for many conductometric biosensors
- The conductivity cell is further miniaturized to provide better reproducibility & minimize the required sample volume.
- Interdigitated microelectrode (IDE) is one good example

Wheatstone Bridge

V_{in}

R_1 R_4

I_a V_{out} I_b

a   b

R_2 R_3

V_{in}

R_1 R_c

V_{out}

R_2 R_3

Conductivity Brid

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Questions and Discussion?