UNIT8-DAY5-LaB1230

Wednesday, April 24, 2013
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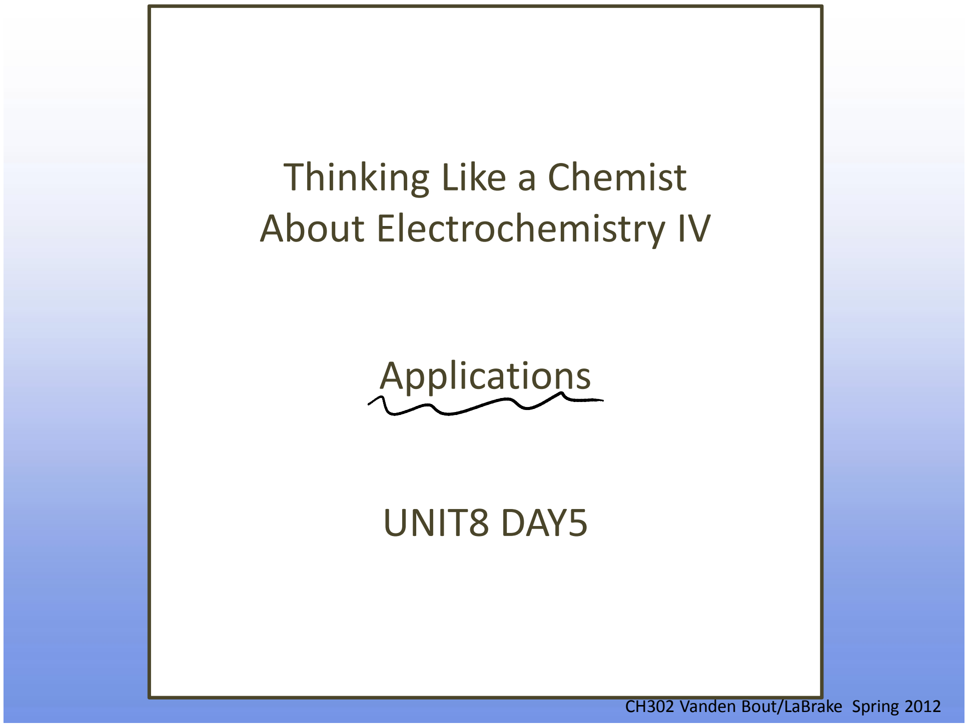
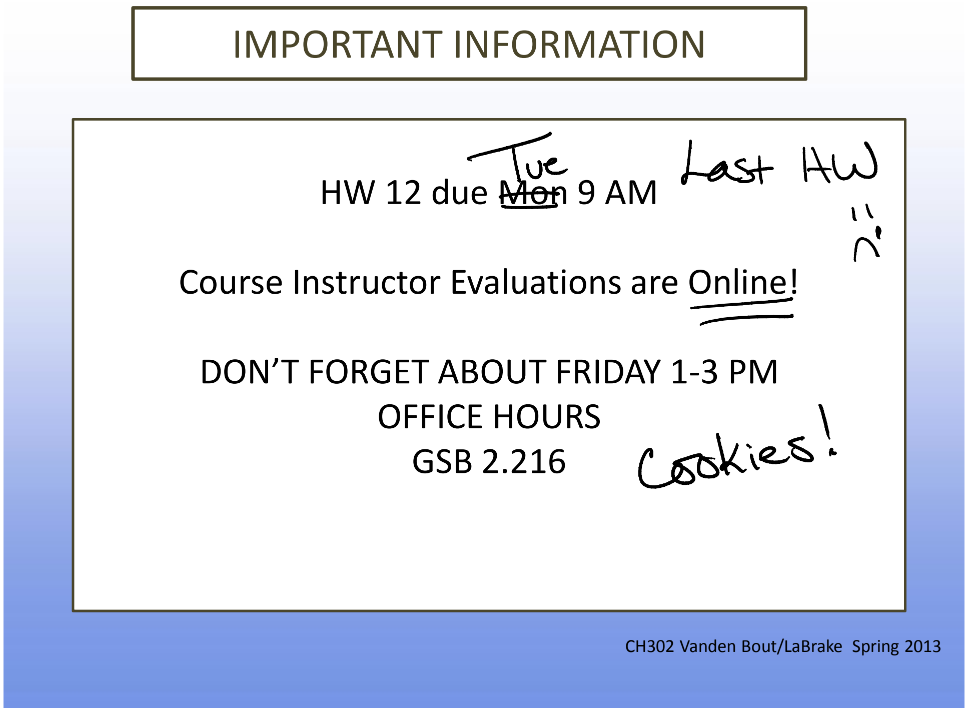


Thinking Like a Chemist
About Electrochemistry IV

Applications

UNIT8 DAY5

CH302 Vanden Bout/LaBrake Spring 2012



IMPORTANT INFORMATION

HW 12 due ~~Mon~~ Tue 9 AM Last HW ☹

Course Instructor Evaluations are Online!

DON'T FORGET ABOUT FRIDAY 1-3 PM
OFFICE HOURS
GSB 2.216 Cookies!

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What are we going to learn today?

Applications of Electrochemistry:

Electrolytic Cells &
Membrane Potential

Quiz: Clicker Question 1

What is the voltage for the following reaction at equilibrium?

dead battery

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

A. 1.1 V → std conditions
[✓ B.] Zero
C. -1.1 V
D. something between 0 and 1.1 V

Quiz: Clicker Question 2

The reaction taking place in a nicad cell is:

$2H_2O$ + Cd(s) + ~~2Ni(OH)~~$_3$(s) → $Cd(OH)_2$(s) + $2Ni(OH)_2$(s)

2Ni(OH)

The emf of the cell when fully charged is 1.25V. = E

What is the reaction free energy? F = 96,485 C/mol e^-

$$\Delta G = -\underbrace{nF}_{\text{charge}}E$$

↳ potential

n = 2

- ~~a) 241 kJ/mol~~
- [✓ b)] -241 kJ/mol
- ~~c) 171 kJ/mol~~
- d) -171 kJ/mol

REVIEW FROM LAST CLASS

Relationship between E and ΔG

ΔG is energy

E is electrical potential

Electric work (energy) = charge X potential

work = -charge X E

$\Delta G = \text{work}_{max}$ $\Delta G = -\text{ charge X } E_{max}$

From now on we'll know the Potential we calculate is the theoretical maximum

Real world never actually that good

→ $\Delta G = -nFE$ ←

Quiz: Clicker Question 3

Assume start at standard conditions

What will happen to the voltage if I lower the Zn^{2+} concentration?

decrease products

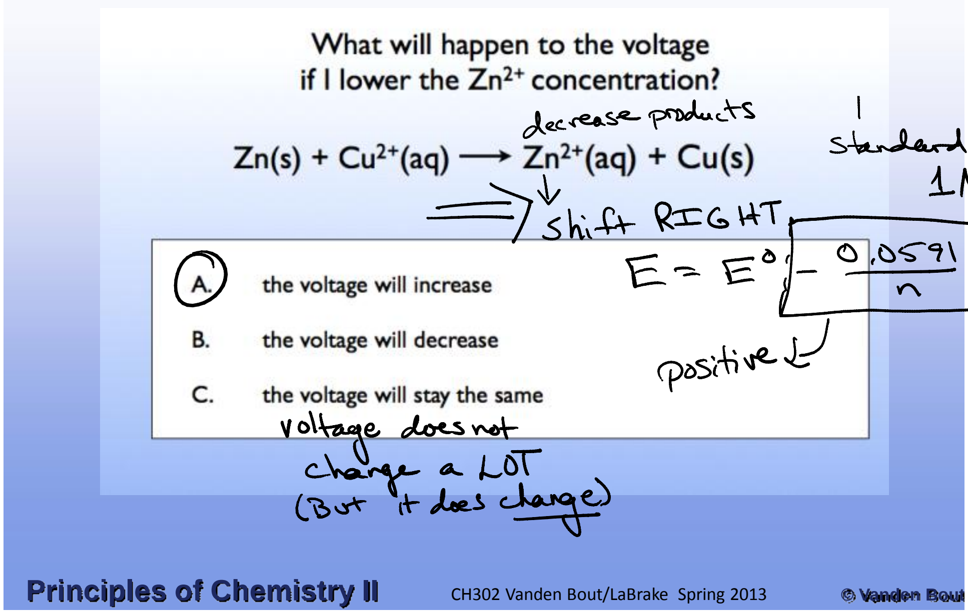
$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

E° = Cathode - anode

↑ standard conditions

1M concentrations

A the voltage will increase



What will happen to the voltage
if I lower the Zn^{2+} concentration?

decrease products

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

$\Longrightarrow$ shift RIGHT

standard conditions
1M concentrations

A. the voltage will increase [✓]

B. the voltage will decrease

C. the voltage will stay the same

voltage does not
change a LOT
(But it does change)

$$\boxed{E = E^0 - \frac{0.0591}{n}\log Q}$$

positive

$$Q = \frac{[Products]}{[Reactants]} = \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$= \frac{0.01}{1}$$

$$\log(1\times10^{-2}) = -2$$

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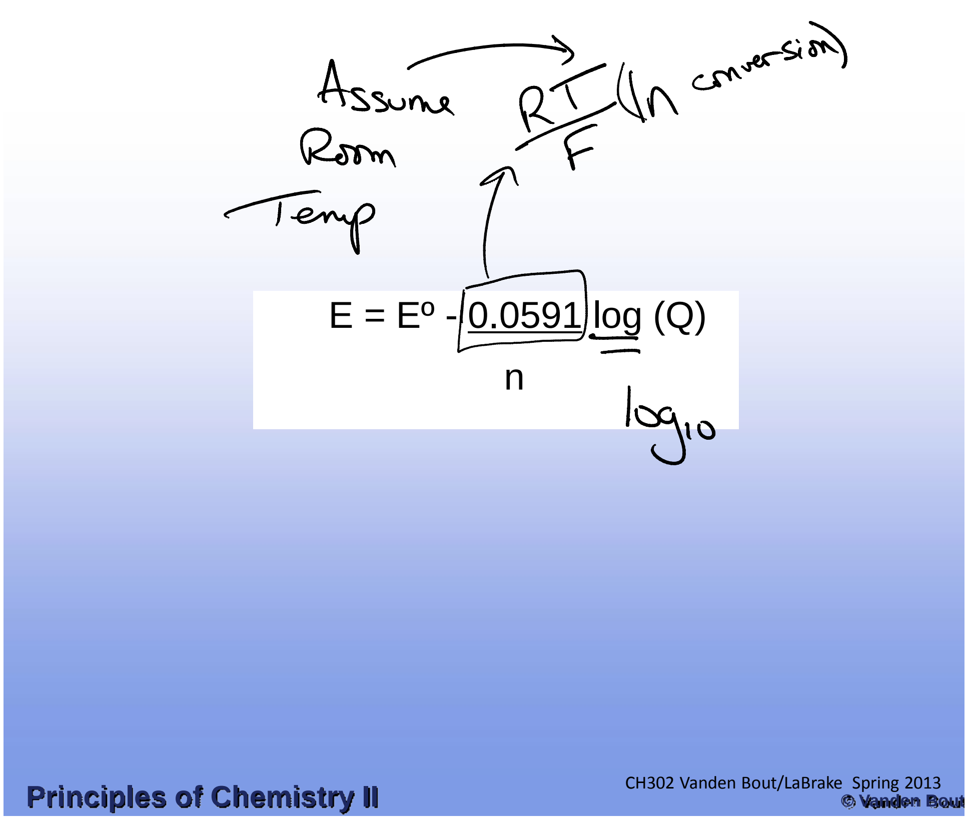

Assume
Room
Temp

$\frac{RT}{F}$ (ln conversion)

$$E = E^o - \frac{\boxed{0.0591}}{n}\log (Q)$$

$\log_{10}$

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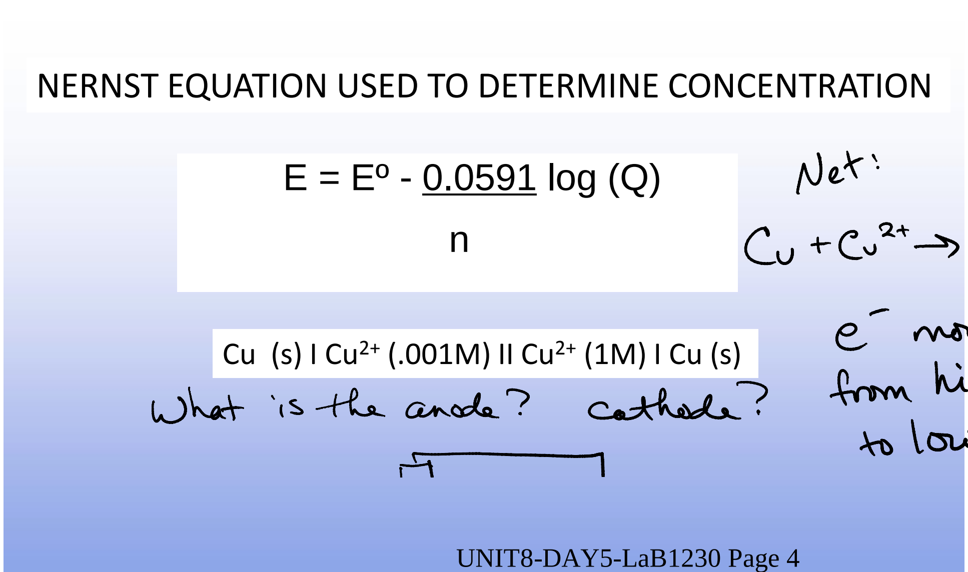


NERNST EQUATION USED TO DETERMINE CONCENTRATION

$$E = E^o - \frac{0.0591}{n}\log (Q)$$

Net:

$Cu + Cu^{2+} \rightarrow Cu^{2+} + Cu$

e^- move
from higher
to lower

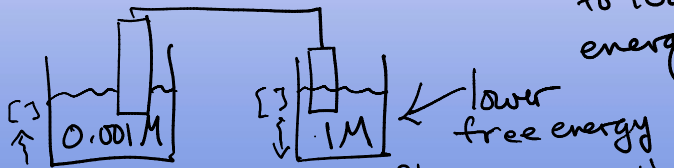
Cu (s) I Cu^{2+} (.001M) II Cu^{2+} (1M) I Cu (s)

What is the anode? cathode?

$Cu\ (s)\ I\ Cu^{2+}\ (.001M)\ II\ Cu^{2+}\ (1M)\ I\ Cu\ (s)$

What is the anode? cathode?

from higher
to lower
energy



$Cu(s) \rightarrow Cu^{2+} + 2e^-$

$Cu^{2+} + 2e^- \rightarrow Cu(s)$

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Poll: Clicker Question 4

Nernst Equation – Dependence of Cell Potential on Concentration

Std Reduction
Potential

$Cu^{2+} + 2e^- \rightarrow Cu(s)$

0.34V

Calculate the emf of the cell:

$Cu(s)|Cu^{2+}(aq, .001\ M)\ ||Cu^{2+}(aq,1\ M)\ |Cu(s)$

a) +0.176 V
b) -0.176 V
[✓ c)] +0.088 V
d) -0.088 V

$$E = E^\circ - \frac{0.0591}{n} \log Q \quad n=2$$

$$E^\circ = 0 \text{ (same concentrations)}$$

$$Q = \frac{0.001}{1} = 1\times10^{-3}$$

$$\log(1\times10^{-3}) = -3$$

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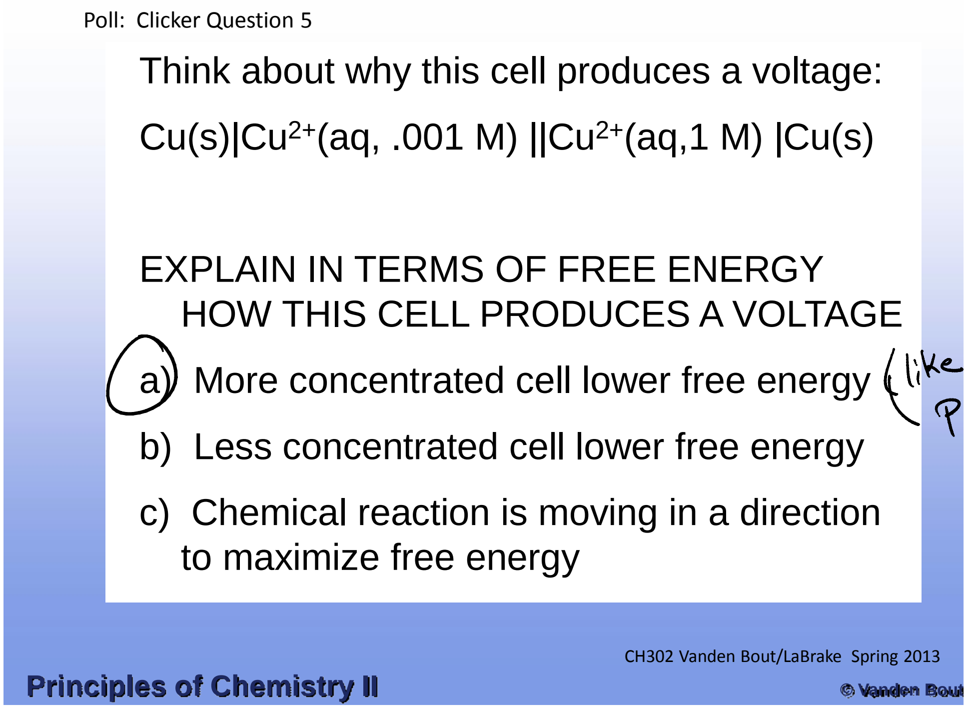
Poll: Clicker Question 5



Think about why this cell produces a voltage:

Cu(s)|Cu^{2+}(aq, .001 M) ||Cu^{2+}(aq,1 M) |Cu(s)

EXPLAIN IN TERMS OF FREE ENERGY HOW THIS CELL PRODUCES A VOLTAGE

[✓ a)] More concentrated cell lower free energy (like vapor pressure!)

b) Less concentrated cell lower free energy

c) Chemical reaction is moving in a direction to maximize free energy

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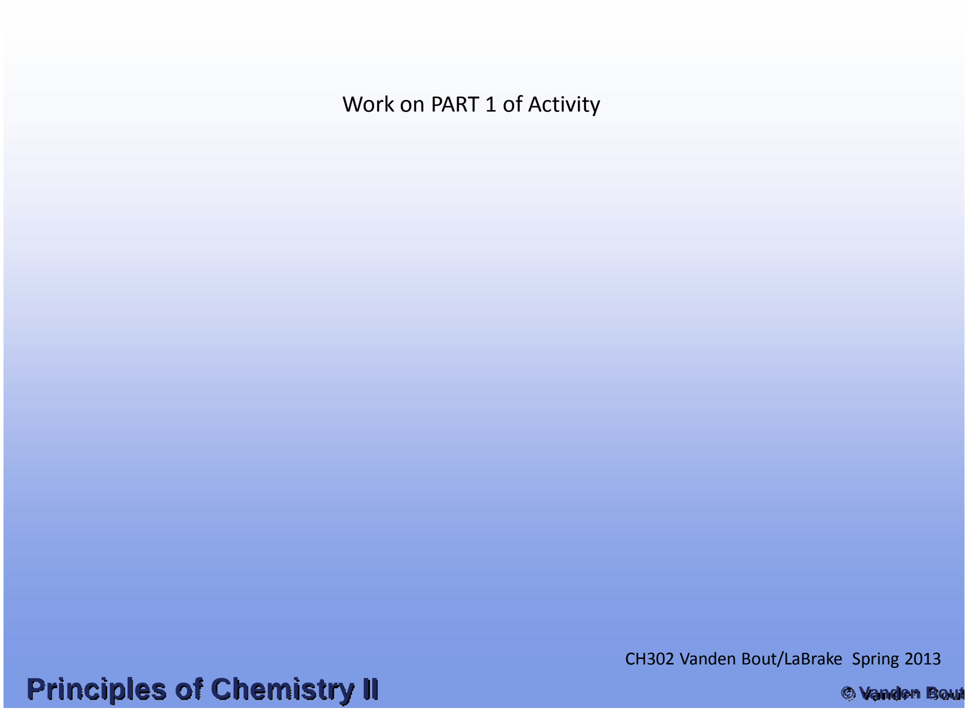

Work on PART 1 of Activity

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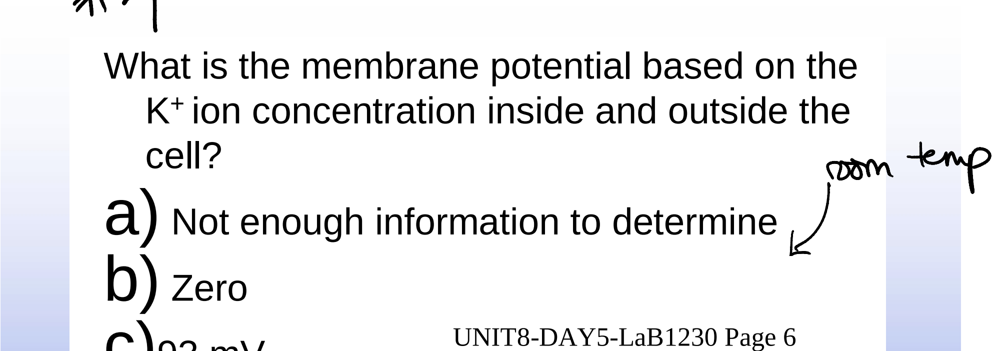
#4



What is the membrane potential based on the K^+ ion concentration inside and outside the cell?

a) Not enough information to determine ← room temp

b) Zero

c) 92 mV

What is the membrane potential based on the K^+ ion concentration inside and outside the cell?

- a) Not enough information to determine
- b) Zero
- [✓ c)] 93 mV
- d) 188 mV
- e) 1 V

room

$$E = \cancel{E}^{0} \searrow 0 - \frac{0.0591}{n} \log Q$$

$n = 1$

$Q = \frac{4}{150}$

Humans ~96mV

Two "kinds" of electrochemical cells

Galvanic (Voltaic): Reaction is spontaneous we can use these to make a battery

create positive voltage

Electrolytic: Reaction is not spontaneous we have to input work to get these reactions to proceed

make it go!
attach external power source

$$2Na(s) + 2H_2O \longrightarrow 2NaOH + H_2(g)$$

What about the reverse?

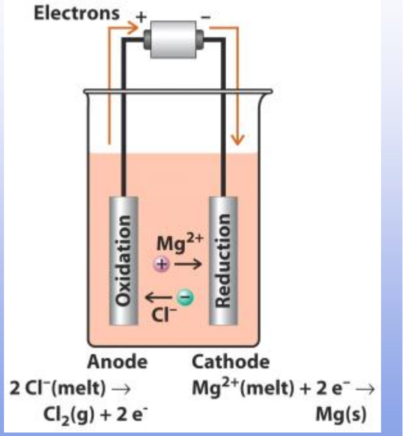
DEMONSTRATE THE REACTIVITY OF SODIUM METAL

HOW DOES ONE GET SODIUM METAL?

http://www.youtube.com/watch?v=i9xS9t-KMpc

Electrolytic Cells

- Reaction is driven in a nonspontaneous direction by using an electric current.
- Anode is site of oxidation but labeled with “+”, and cathode is site of reduction but labeled with a “-”.
- Over potential is the extra emf over the emf of the cell that is needed to push the reaction in a nonspontaneous direction.



Example of Electrolytic Cell

Consider the electrolysis of aqueous copper(II)bromide. The observations are: one electrode becomes coated with copper metal, and the color of the solution around the electrode fades; around the other electrode the solution turns brown, as the bromine is formed and some O_2 bubbles are formed.

Design the cell, label electrodes, flow of e^-, and ½ reactions.

DO PART II OF ACTIVITY

What did we learn today?

Recognize and explain applications of electrochemistry including batteries, electrolytic cells, and cell membranes.