EXPERIMENT 10

Redox Reactions & Voltaic/Galvanic Cells

DR. MIOY T. HUYNH
CHEMISTRY 136L
YALE UNIVERSITY
FALL 2018

SPONTANEOUS REACTIONS

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

This reaction is spontaneous.

How do we know this?

When you bring together Zn metal and aqueous Cu^{2+}, there will be a spontaneous change.

Zn can reduce Cu^{2+} to Cu metal.
This implies Cu is unable to reduce Zn^{2+}.
In other words, the reverse reaction is *nonspontaneous*.

Zn metal is a stronger reducing agent than Cu metal.

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

Cu metal can reduce Ag^{+} to Ag metal.

Cu metal is a stronger reducing agent than Ag metal.

Zn is the strongest reductant.
Ag is the weakest reductant.

VOLTAIC/GALVANIC CELL

Zn Electrode: negative electrode → OXIDATION

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

Cu Electrode: positive electrode → REDUCTION

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

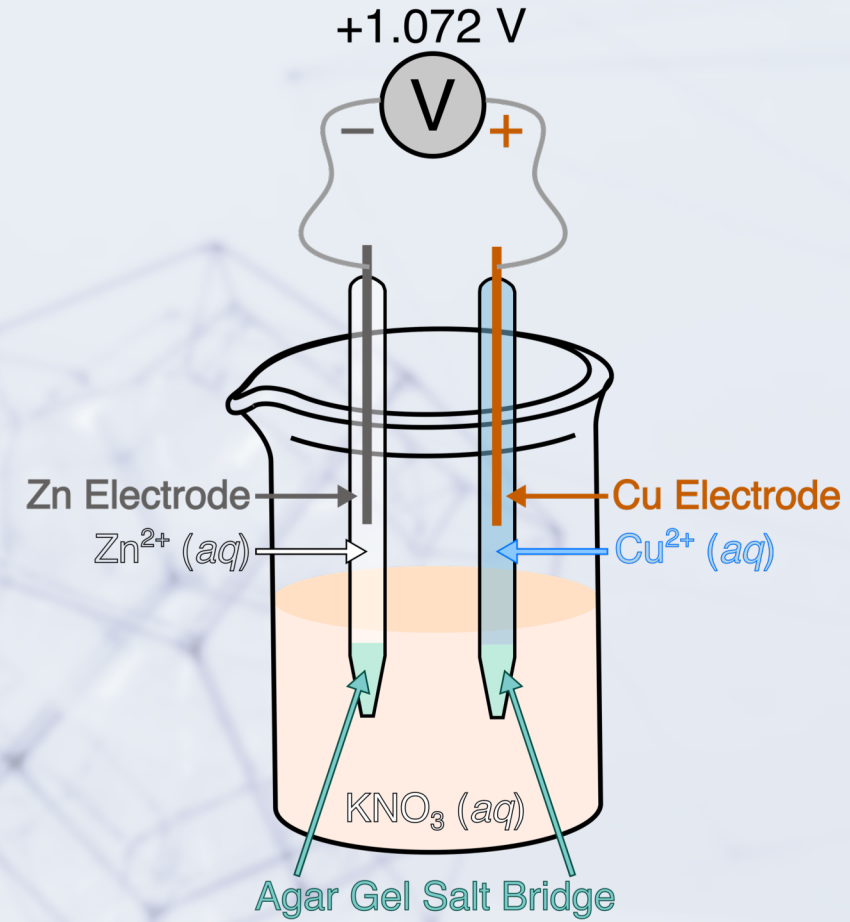
The ELECTRONS released from OXIDATION of Zn metal makes the Zn electrode the negative electrode.

↓

These ELECTRONS leave the negative Zn electrode and flow through the external circuit (a VOLTMETER).

↓

When these ELECTRONS arrive at the positive Cu electrode, Cu^{2+} ions are REDUCED to Cu metal.



The opposite is true in an ELECTROLYTIC CELL where a *nonspontaneous* reaction is forced to occur by electrical energy input.

OXIDATION now takes place at the positive electrode.
REDUCTION now takes place at the negative electrode.

Since the overall reaction is the sum of the two half-reactions, we have:

$$EMF_{cell} = EMF_{Zn\ half\text{-}cell} + EMF_{Cu\ half\text{-}cell}$$

$$E_{cell} = E_{Zn} + E_{Cu}$$

EMF stands for electromotive forces, measured in Volts (V).

VOLTAIC/GALVANIC CELL $\Delta G = -nFE_{cell}$

n = number of moles of electrons in overall reaction
F = Faraday constant (charge, C, on one mole of electrons)
E_{cell} = EMF of the cell

SPONTANEOUS REACTION $\rightarrow \Delta G < 0 \rightarrow E_{cell} > 0$
NONSPONTANEOUS REACTION $\rightarrow \Delta G > 0 \rightarrow E_{cell} < 0$

VOLTAIC/GALVANIC CELL

Zn Electrode: negative electrode → OXIDATION

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

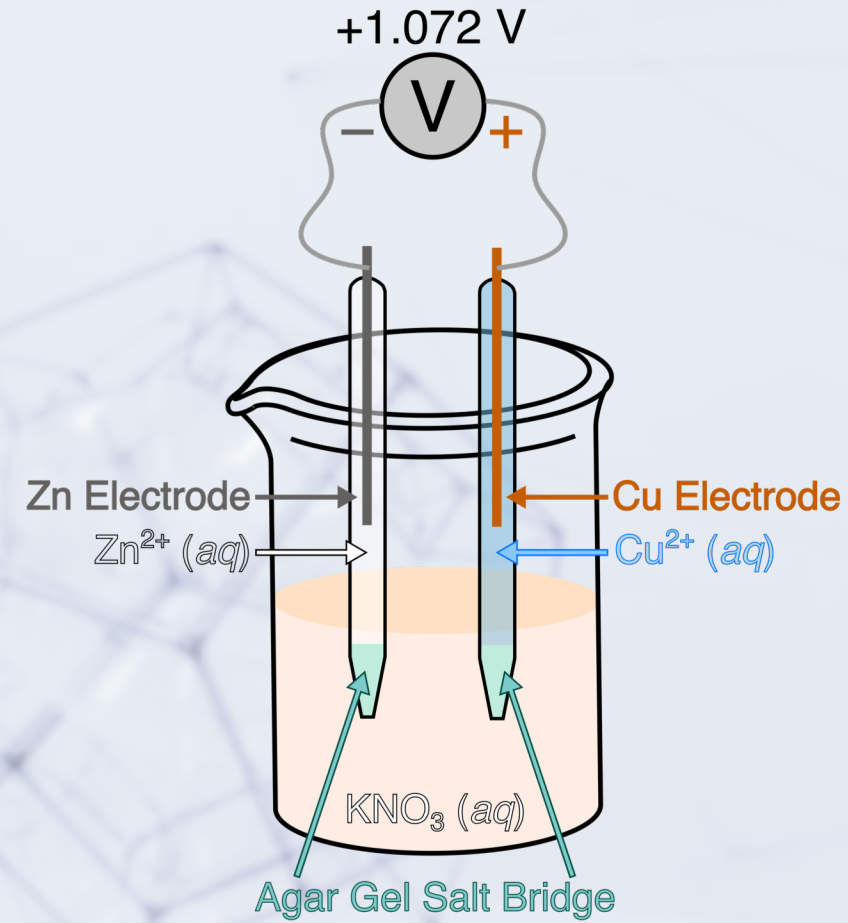
Cu Electrode: positive electrode → REDUCTION

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

We can measure the E_{cell} using a voltmeter.

Can we assign individual half-cell *EMF*s?

YES!



CONVENTION
Standard Hydrogen Electrode (SHE)

$H_2\,(g) \rightarrow 2\,H^+\,(aq) + 2\,e^-$; $E° = 0.0$ V

Combine the Cu half-cell with the H_2 half-cell (SHE).

↓

The measured E_{cell} is then:
+0.337 V with SHE as the negative electrode

↓

Therefore, we know:
$Cu^{2+}\,(aq) + 2\,e^- \rightarrow Cu\,(s)$; $E° = +0.337$ V

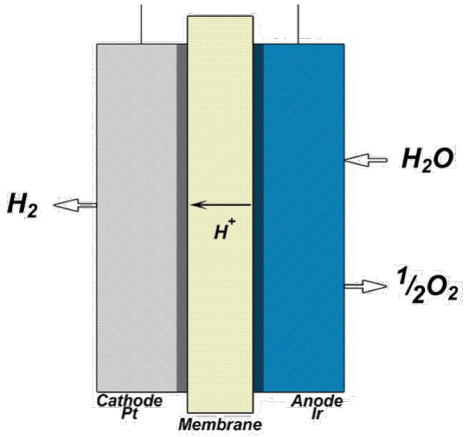
The H_2/H^+ half-cell is the PRIMARY REFERENCE HALF-CELL.
The Cu/Cu^{2+} half-cell is the SECONDARY REFERENCE HALF-CELL.

SOLAR ELECTROLYSIS OF WATER

POSITIVE electrode $H_2O\ (\ell) \rightarrow 2\ H^+\ (aq) + \frac{1}{2}\ O_2\ (g) + 2\ e^-$

NEGATIVE electrode $2\ H^+\ (aq) + 2\ e^- \rightarrow H_2\ (g)$

H^+ migrates across a PROTON EXCHANGE MEMBRANE (PME) to the other side.



H_2—O_2 FUEL CELL

POSITIVE electrode $2\ H^+\ (aq) + \frac{1}{2}\ O_2\ (g) + 2\ e^- \rightarrow H_2O\ (g)$

NEGATIVE electrode $H_2\ (g) \rightarrow 2\ H^+\ (aq) + 2\ e^-$

H_2 and O_2 gas combine to form H_2O spontaneously.
This is a voltaic/Galvanic cell.

NOTES

1. The graphite electrodes can break easily.
   Attach it to the alligator clip first, then place into gel tube.

2. You have to exchange your ID for the platinum electrode.

3. Next week:
   - Make up any labs you missed
   - Check out