Electrochemistry

Harnessing the changes in oxidation and reduction

Zn^{2+} + Cu^{2+}

(aq) (s) (aq) (s)

Remember:

- Anode: electrode in the half-cell where oxidation takes place
- Metal electrode atoms are oxidized and become aqueous ions
- Anions must flow from the salt bridge into this half cell to balance out the addition of new metal ions

Remember:

- Cathode: electrode in the half-cell where reduction takes place
- Metal ions become neutral atoms
- Cations must flow from the salt bridge into this half cell to balance loss of metal ions

Remember:

- Electrons flow through the wire from the anode towards the cathode
- Anions flow from the salt bridge into the anode half-cell
- Cations flow from the salt bridge into the cathode half-cell

A few terms…

- The flow of electrons through the wire is referred to as “current”
- Electric current is measured in amperes (or “amps”)
- If either the wire or salt bridge are removed, current ceases to flow

A few terms…

- Potential = the force exerted on the electrons in a wire or other conductor causing them to flow
- Measured as "volts"; often referred to as "voltage"

∴ <u>voltage</u> is a measure of <u>force</u>

Reduction potential:

- The potential ("likelihood") for a half cell to undergo reduction
- Recall: *reduction = gain (take)* e^-*'s*
- *typically*: metal ions (M^+) are reduced to metal atoms (M^o)
- Measured as "volts"
    - The measure of the **pull** on the electrons

- When two half-cells are connected by a wire and salt bridge, *the half-cell with the greater reduction potential gets* ***reduced***
  - It “wins” the electrons
- The other half-cell gets *oxidized*
  - It “loses” the electrons

Cell potential

The cell potential (E_{cell}) is the difference between the two reduction potentials of the two half cells

$$E_{cell} = E_{red} - E_{ox}$$

For a reaction to happen, the E_{cell} must be a **positive number**

“Standard” Cell potential

- The E°_{cell} is the “standard” cell potential
- That means
  - –All solutions are 1.0M
  - –Pressure = 1.0atm
  - –(Temperature = 25°C)

Cell potential

- All E°'s for half cells are arbitrary numbers
  - They are based on deciding the "standard hydrogen electrode" has an E° = 0.0V

- All other reduction potentials are measured relative to this value

Cell notation

anode | anode ion || cathode ion | cathode

Ex: Zn | Zn^{2+} || Cu^{2+} | Cu

Think: half reactions

$$Zn \rightarrow Zn^{2+} + 2\ e^-$$

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

Cell notation

anode | anode ion || cathode ion | cathode

Ex: Zn | Zn^{2+} || Cu^{2+} | Cu

$E°_{cell} = E°_{Cu} - E°_{Zn}$

$E°_{cell} = 0.34V – (-0.76V)$

$E°_{cell} = 1.10V$

anode | anode ion || cathode ion | cathode

What is the cell notation for:

$$Mg_{(s)} + Fe^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Fe_{(s)}$$

$$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^{-}$$ (ox = anode)

$$Fe^{2+}_{(aq)} + 2e^{-} \rightarrow Fe_{(s)}$$ (red = cathode)

$$Mg \,|\, Mg^{2+} \,||\, Fe^{2+} \,|\, Fe$$

Cell notation

What is $E°_{cell}$ for the reaction?

$Mg \mid Mg^{2+} \parallel Fe^{2+} \mid Fe$

$E°_{cell} = E°_{Fe} - E°_{Mg}$

$E°_{cell} = -0.44V – (-2.37V)$

$E°_{cell} = 1.93V$

Cell notation

What is the reaction for the cell notation

$$Al \mid Al^{3+} \mid\mid Pb^{2+} \mid Pb$$

Al is being oxidized: $Al \rightarrow Al^{3+} + 3e^-$

Pb^{2+} is being reduced: $Pb^{2+} + 2e^- \rightarrow Pb$

- Multiply through to “balance the electrons”:

$$2\ Al + 3\ Pb^{2+} \rightarrow 2\ Al^{3+} + 3\ Pb$$

Cell notation

What is $E°_{cell}$ for the reaction?

$Al \mid Al^{3+} \parallel Pb^{2+} \mid Pb$

$E°_{cell} = E°_{Pb} - E°_{Al}$

$E°_{cell} = -0.13V – (-1.66V)$

$E°_{cell} = 1.53V$

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(aq) (s) (aq) (s)