

Electrochemistry

Harnessing the changes
in oxidation and
reduction

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”
- ∴ voltage is a measure of force

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^0)
- 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

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

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 $E^\circ = 0.0V$
- All other reduction potentials are measured relative to this value

Cell notation

anode | anode ion || cathode ion | cathode

Ex: $\text{Zn} \mid \text{Zn}^{2+} \parallel \text{Cu}^{2+} \mid \text{Cu}$

Think: half reactions



Cell notation

anode | anode ion || cathode ion | cathode

Ex: Zn | Zn²⁺ || Cu²⁺ | Cu

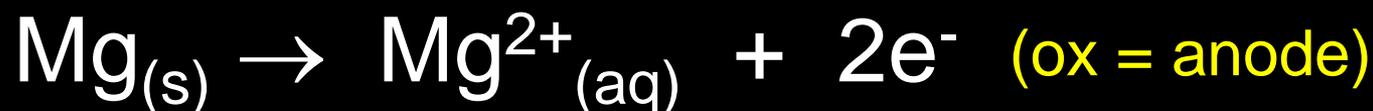
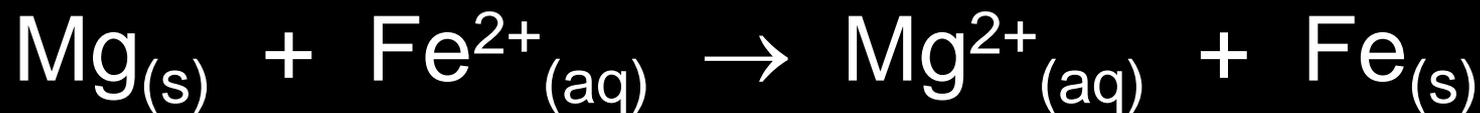
$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Cu}} - E^{\circ}_{\text{Zn}}$$

$$E^{\circ}_{\text{cell}} = 0.34\text{V} - (-0.76\text{V})$$

$$E^{\circ}_{\text{cell}} = 1.10\text{V}$$

anode | anode ion || cathode ion | cathode

What is the cell notation for:



Cell notation

What is E°_{cell} for the reaction?



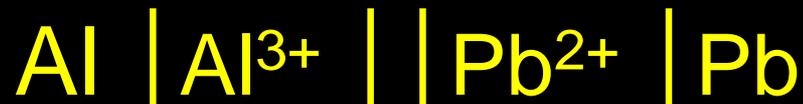
$$E^\circ_{\text{cell}} = E^\circ_{\text{Fe}} - E^\circ_{\text{Mg}}$$

$$E^\circ_{\text{cell}} = -0.44\text{V} - (-2.37\text{V})$$

$$E^\circ_{\text{cell}} = 1.93\text{V}$$

Cell notation

What is the reaction for the cell notation



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

Pb²⁺ is being reduced: $\text{Pb}^{2+} + 2\text{e}^{-} \rightarrow \text{Pb}$

- Multiply through to “balance the electrons”:



Cell notation

What is E°_{cell} for the reaction?



$$E^\circ_{\text{cell}} = E^\circ_{\text{Pb}} - E^\circ_{\text{Al}}$$

$$E^\circ_{\text{cell}} = -0.13\text{V} - (-1.66\text{V})$$

$$E^\circ_{\text{cell}} = 1.53\text{V}$$

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