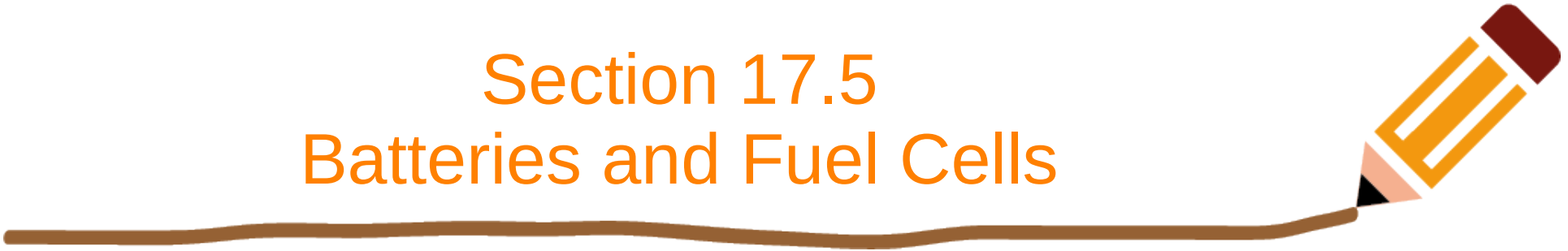
Section 17.5
Batteries and Fuel Cells



Michael Stogsdill

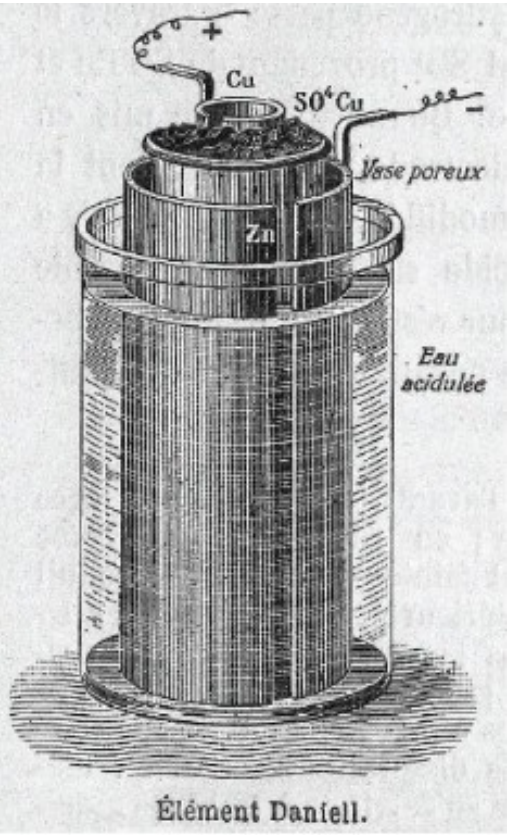
Mott Community College

Learning Objectives

- Describe the electrochemistry associated with several common batteries
- Distinguish the operation of a fuel cell from that of a battery

Batteries

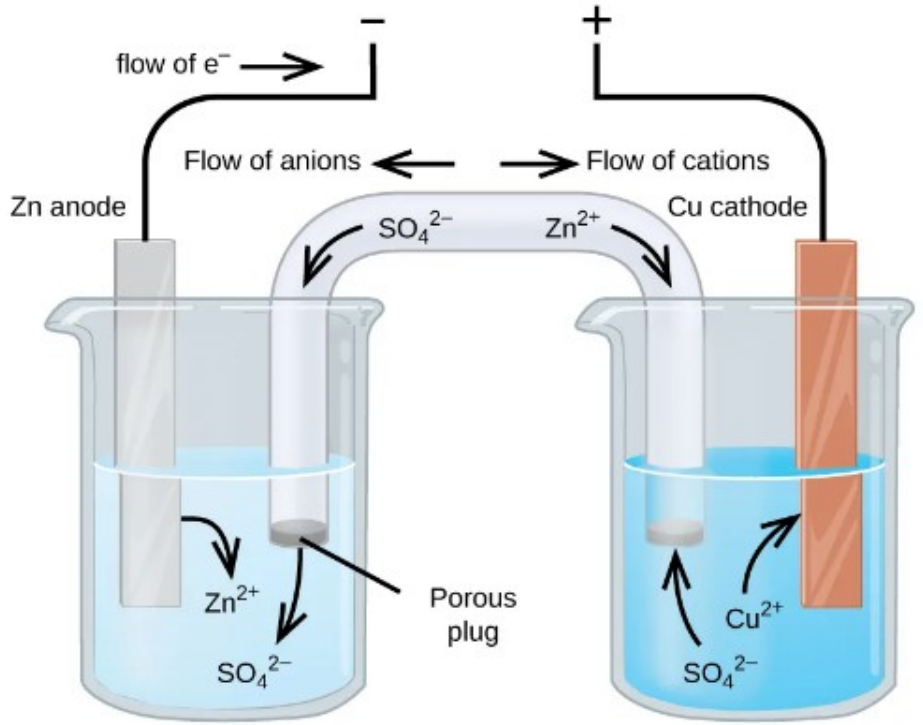
- A **battery** is a galvanic cell that has been specially designed and constructed in a way that best suits its intended use a source of electrical power for specific applications.
- Among the first successful batteries was the *Daniell cell*, which relied on the spontaneous oxidation of zinc by copper(II) ions.



Élément Daniell.

Daniell cell

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

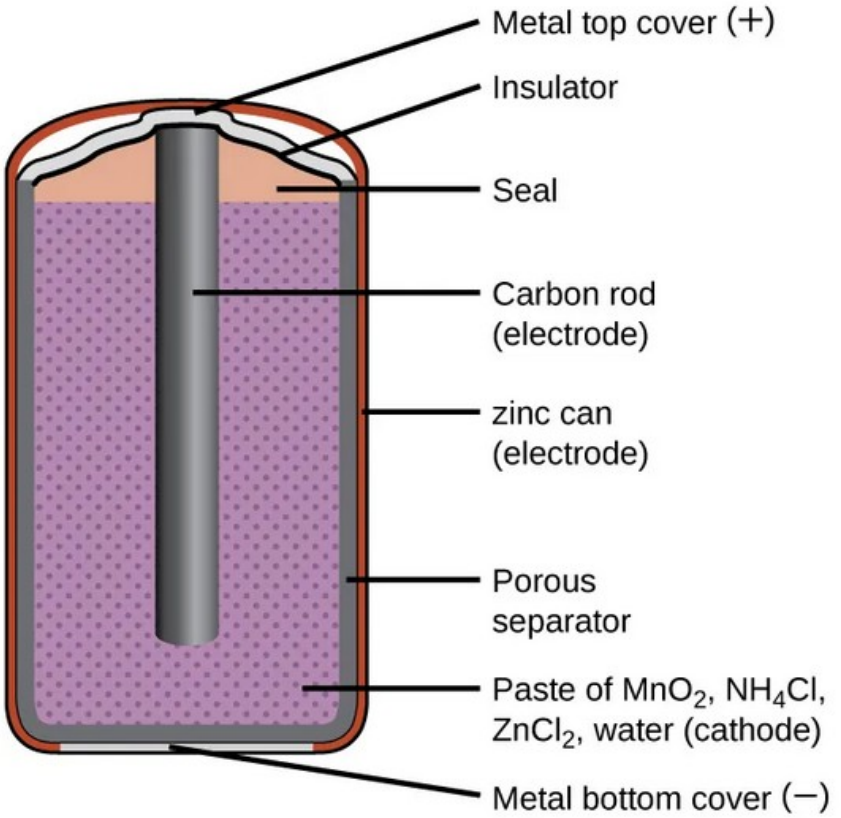


Modern Batteries

- Modern batteries exist in a multitude of forms to accommodate various applications.
  - Tiny button batteries that provide the power needs of a wristwatch
  - Very large batteries used to supply backup energy to municipal power grids.
- Some batteries are designed for single-use applications and cannot be recharged (**primary cells**).
- Others are based on conveniently reversible cell reactions that allow recharging (**secondary cells**).

Dry Cells

- A common primary battery is the **dry cell**.
  - zinc is used as both container and anode (“–” terminal)
  - A graphite rod as the cathode (“+” terminal)
  - An electrolyte paste containing manganese(IV) oxide, zinc(II) chloride, ammonium chloride, and water bridges the two electrodes.



Dry Cell Voltage

- Zinc is oxidized at the anode.

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

- Manganese (IV) is reduced within the paste.

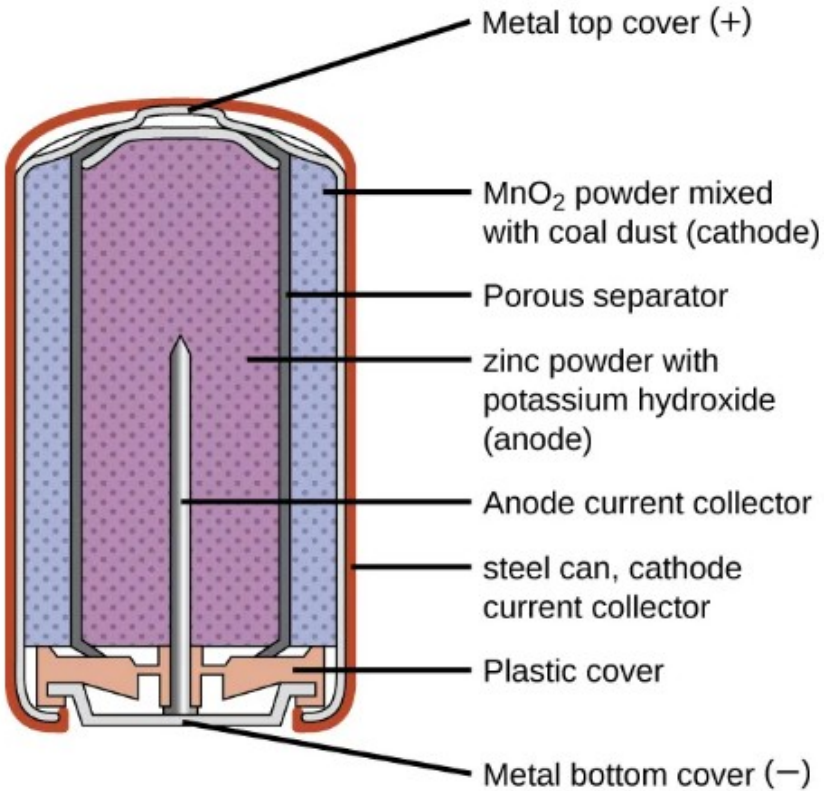
$$2\mathrm{MnO_2}(s) + 2\mathrm{NH_4Cl}(aq) + 2\,\mathrm{e}^- \longrightarrow \mathrm{Mn_2O_3}(s) + 2\mathrm{NH_3}(aq) + \mathrm{H_2O}(l) + 2\mathrm{Cl}^-$$

- The net reaction has a cell potential of 1.5 V.

$$2\mathrm{MnO_2}(s) + 2\mathrm{NH_4Cl}(aq) + \mathrm{Zn}(s) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Mn_2O_3}(s) + 2\mathrm{NH_3}(aq) + \mathrm{H_2O}(l) + 2\mathrm{Cl}^-$$

Alkaline batteries

- **Alkaline batteries** were developed in the 1950s to improve on the performance of the dry cell
- As their name suggests, these types of batteries use alkaline electrolytes
  - Often potassium hydroxide.
- An alkaline battery can deliver about three to five times the energy of a zinc-carbon dry cell of similar size.

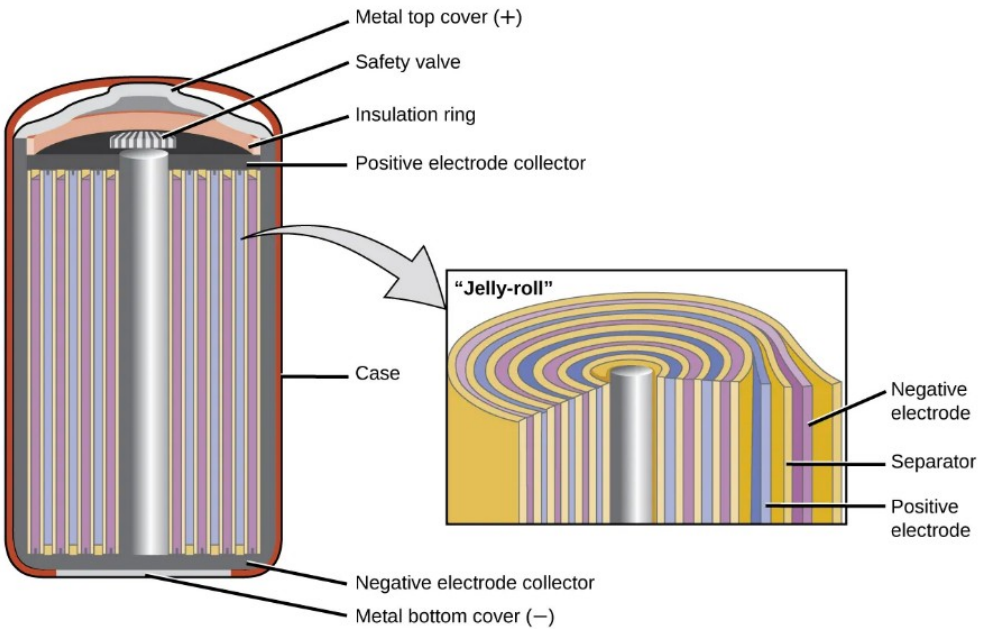


Alkaline Cell Voltage

$$\begin{array}{lrcll}
\text{anode:} & \mathrm{Zn}(s) + 2\mathrm{OH}^-(aq) & \longrightarrow & \mathrm{ZnO}(s) + \mathrm{H_2O}(l) + 2\mathrm{e}^- & \\
\text{cathode:} & 2\mathrm{MnO_2}(s) + \mathrm{H_2O}(l) + 2\mathrm{e}^- & \longrightarrow & \mathrm{Mn_2O_3}(s) + 2\mathrm{OH}^-(aq) & \\
\hline
\text{cell:} & \mathrm{Zn}(s) + 2\mathrm{MnO_2}(s) & \longrightarrow & \mathrm{ZnO}(s) + \mathrm{Mn_2O_3}(s) & E_{\text{cell}} = +1.43\ \text{V}
\end{array}$$

Nickel-Cadmium Cells

- **Nickel-cadmium, or NiCd**, batteries consist of a nickel-plated cathode, cadmium-plated anode, and a potassium hydroxide electrode.
  - A “jelly-roll” design and allows it to deliver more current than a similar-sized alkaline battery.
  - A NiCd battery can be recharged about 1000 times.
  - Cadmium is a toxic heavy metal



Nickel-Cadmium Cell Voltage

$$\text{anode:} \quad \mathrm{Cd}(s) + 2\mathrm{OH}^-(aq) \longrightarrow \mathrm{Cd(OH)_2}(s) + 2\mathrm{e}^-$$

$$\text{cathode:} \quad \mathrm{NiO_2}(s) + 2\mathrm{H_2O}(l) + 2\mathrm{e}^- \longrightarrow \mathrm{Ni(OH)_2}(s) + 2\mathrm{OH}^-(aq)$$

$$\text{cell:} \quad \mathrm{Cd}(s) + \mathrm{NiO_2}(s) + 2\mathrm{H_2O}(l) \longrightarrow \mathrm{Cd(OH)_2}(s) + \mathrm{Ni(OH)_2}(s) \qquad E_{\text{cell}} \text{-} 1.2\ \mathrm{V}$$

Lithium Ion Cells

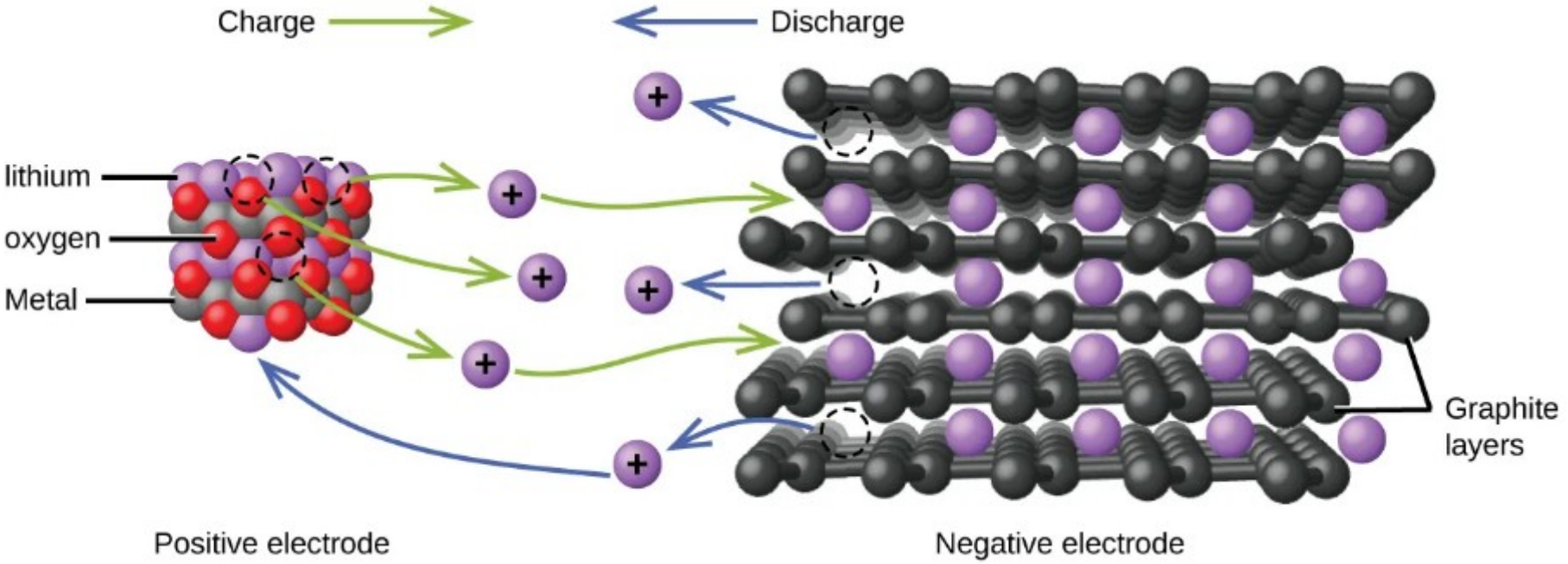
- **Lithium ion batteries** are among the most popular rechargeable batteries and are used in many portable electronic devices.
  - Provide a large amount current
  - Are lighter than comparable batteries of other types
  - Produce a nearly constant voltage as they discharge
  - Only slowly lose their charge when stored.

Lithium Cell Voltage

- The variable stoichiometry of the cell reaction leads to variation in cell voltages
  - *x* is usually no more than 0.5 and the cell voltage is 3.7 V.

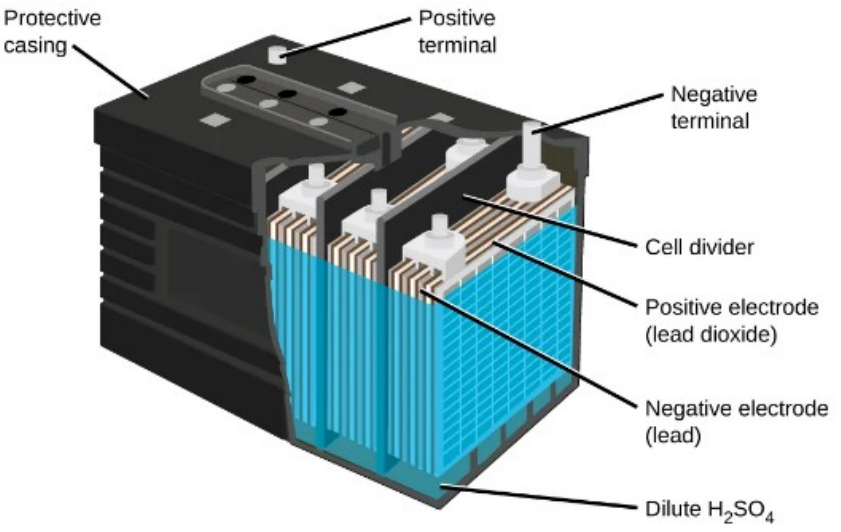
$$\begin{aligned}
&\text{anode:} && \mathrm{LiCoO_2} \rightleftharpoons \mathrm{Li_{1-x}CoO_2} + x\,\mathrm{Li^+} + x\,\mathrm{e^-} \\
&\text{cathode:} && x\,\mathrm{Li^+} + x\,\mathrm{e^-} + x\,\mathrm{C_6} \rightleftharpoons x\,\mathrm{LiC_6} \\
\hline
&\text{cell:} && \mathrm{LiCoO_2} + x\,\mathrm{C_6} \rightleftharpoons \mathrm{Li_{1-x}CoO_2} + x\,\mathrm{LiC_6} \qquad E_{\text{cell}} \text{-} 3.7\,\mathrm{V}
\end{aligned}$$

Lithium Ion Transfer



Lead Acid Batteries

- The lead acid battery is the type of secondary battery commonly used in automobiles.
  - It is inexpensive and capable of producing high current.
  - Lead acid batteries are heavy and contain a caustic liquid electrolyte, H_2SO_4 (*aq*)
  - They must always be disposed of properly.



Lead Acid Cell Voltage

- Each cell produces 2 V, six cells are connected in series to produce a 12-V car battery.

$$\text{anode:} \quad \mathrm{Pb}(s) + \mathrm{HSO_4}^-(aq) \longrightarrow \mathrm{PbSO_4}(s) + \mathrm{H}^+(aq) + 2\mathrm{e}^-$$

$$\text{cathode:} \quad \mathrm{PbO_2}(s) + \mathrm{HSO_4}^-(aq) + 3\mathrm{H}^+(aq) + 2\mathrm{e}^- \longrightarrow \mathrm{PbSO_4}(s) + 2\mathrm{H_2O}(l)$$

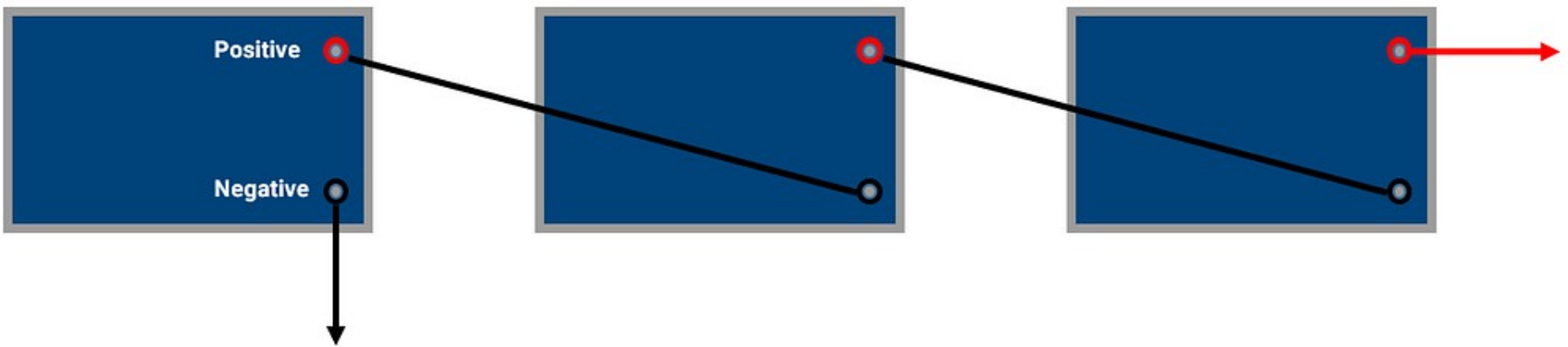
$$\text{cell:} \quad \mathrm{Pb}(s) + \mathrm{PbO_2}(s) + 2\mathrm{H_2SO_4}(aq) \longrightarrow 2\mathrm{PbSO_4}(s) + 2\mathrm{H_2O}(l) \qquad E_{\text{cell}} = 2\ \text{V}$$

Battery Capacity

- Batteries come in different sizes. Sometimes this is combine multiple cells or contain one larger cells.
  - Dry cells are available in various sizes (e.g., D, C, AA, AAA).
- Larger cells contain greater amounts of the redox reactants and therefore are capable of transferring correspondingly greater amounts of charge before being spent.
- Larger cells will not be able to deliver more current or greater voltage.
  - Increasing *E* or *i* requires combining multiple cells.

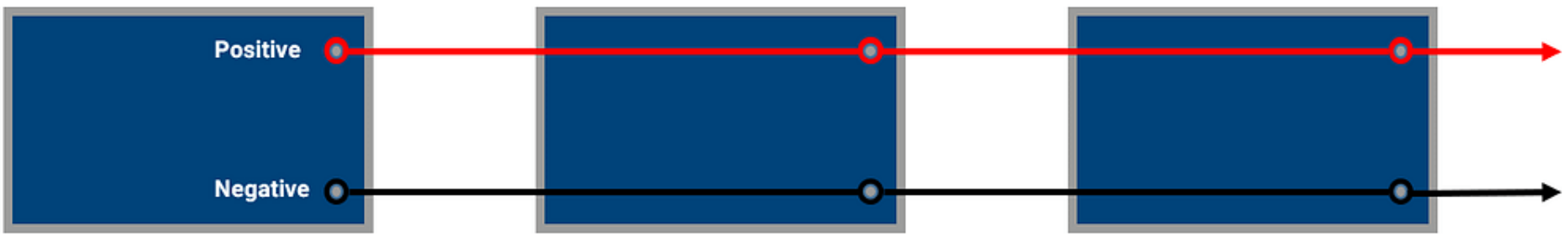
Batteries in Series

- Cells can be combined in series to increase the voltage of the overall battery.
- Each cell in series will increase the batteries voltage by the potential of a cell.
  - Each dry cell is 1.5 V, four dry cells in series would produce 6 V



Batteries in Parallel

- Cells can be combined in parallel to increase the capacity and maximum current available.
- Each cell in parallel will increase the battery capacity by the capacity of a single cell.
  - Wiring three 30 AH lead acid batteries will result in a 90 AH battery pack.



Fuel Cells

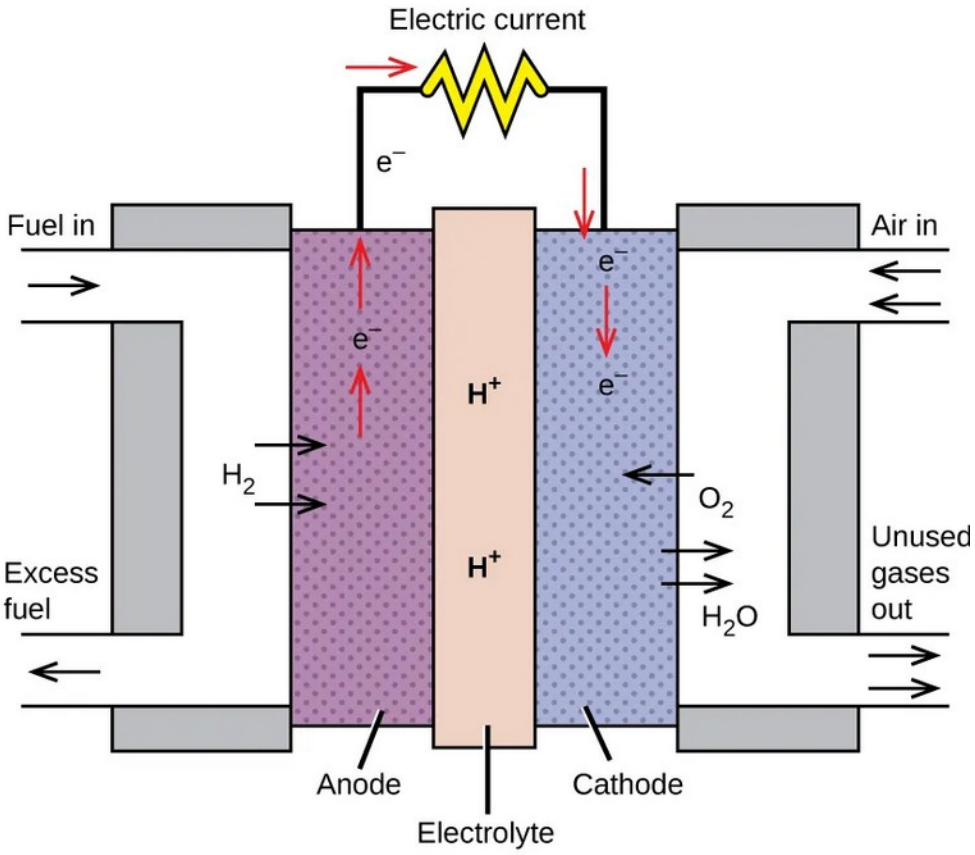
- A **fuel cell** is a galvanic cell that uses traditional combustive fuels that are continuously fed into the cell along with an oxidant.
  - Typical fuels include hydrogen, methane, and methanol
- Within the cell, fuel and oxidant undergo the same redox chemistry as when they are combusted
  - The catalyzed electrochemical process is much more efficient than traditional combustion.

Fuel Cell Example

- A typical hydrogen fuel cell uses graphite electrodes embedded with platinum-based catalysts to accelerate the two half-cell reactions

$$\begin{aligned}
&\text{Anode:} && 2\mathrm{H_2}(g) \longrightarrow 4\mathrm{H^+}(aq) + 4\mathrm{e^-} \\
&\text{Cathode:} && \mathrm{O_2}(g) + 4\mathrm{H^+}(aq) + 4\mathrm{e^-} \longrightarrow 2\mathrm{H_2O}(g) \\
\hline
&\text{Cell:} && 2\mathrm{H_2}(g) + \mathrm{O_2}(g) \longrightarrow 2\mathrm{H_2O}(g) \qquad E_{\text{cell}} \text{-} 1.2\ \mathrm{V}
\end{aligned}$$

Fuel Cell Example



- These types of fuel cells generally produce approximately 1.2 V.
- The energy efficiency is 20%–25% for an engine versus ~50%–75% for a fuel cell