• Apr 24  Ch 7
• Apr 26  Ch 8 Letter due

• May 1  Ch 8
• May 3  Ch 8 Q 10, HW 9

• May 8  Ch 8?
• May 10  **Exam 3** (Ch 5, 7, 8) HW 10

• May 15  Review and Wrap-up
Chapter 8:
Energy From Electron Transfer
Consider the Battery

On the opposite end of the scale from the power plant is the battery
Personal, portable power supply
But what IS a battery?
How does it work?
Why can some be recharged and some can’t?
Are there alternatives to traditional batteries?
Can batteries (or their alternatives) help with the energy crunch?
Electrochemistry: Some Definitions

A Battery: A system which converts chemical energy into electrical energy

More correctly, a battery is an electrochemical cell:

Galvanic Cells convert the energy from spontaneous chemical reactions into electricity

Electrolytic Cells use electricity to drive non-spontaneous chemical reactions
Electrochemistry: Some Definitions

All galvanic cells produce electricity from reactions which involve the transfer of electrons from one species to another.

There are two components to each cell – the species donating the electrons, and the species accepting them.

We write “half-reactions” to represent these two components, and to explicitly show the transfer of electrons.
Electrochemistry: Some Definitions

The **oxidation half-reaction** shows the species which is donating electrons

The **reduction half-reaction** show the species which is receiving electrons

We can also write the **net reaction** (or overall reaction) for the cell, the balanced sum of the two half-reactions

LEO the lion says GER:

Loss of Electrons is Oxidation; Gain of Electrons is Reduction
In a nickel-cadmium battery, the reactions look something like this:

Oxidation \[ \text{Cd} \rightarrow \text{Cd}^{2+} + 2 \text{e}^- \]

Reduction \[ \text{Ni}^{3+} + \text{e}^- \rightarrow \text{Ni}^{2+} \]

Net \[ \text{Cd} + 2 \text{Ni}^{3+} \rightarrow \text{Cd}^{2+} + 2 \text{Ni}^{2+} \]

Note: The number of electrons given off in the oxidation half-reaction must equal the number gained in the reduction half-reaction.

Electrons moving from one place to another – this is electricity.
Electrodes are electrical conductors in the cell where chemical reactions take place.
The anode is the electrode where oxidation takes place.
The cathode is the electrode where reduction takes place.
The cathode receives the electrons given off at the anode and passes them along.
The voltage of the whole cell is the electrical energy that it gives off, measured in volts (V).
The current is the rate at which electrons pass through the cell, measured in amperes (A).
A voltaic cell diagram showing a zinc (anode) and copper (cathode) half-cells connected by a salt bridge. The cell reaction is:

\[ \text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s) \]
In a nickel-cadmium battery, the reactions actually look like this:

**Oxidation**

\[ \text{Cd}(s) + 2 \text{OH}^- (aq) \rightarrow \text{Cd(OH)}_2(s) + 2 \text{e}^- \]

**Reduction**

\[ 2\text{NiO(OH)}(s) + 2 \text{H}_2\text{O(l)} + 2 \text{e}^- \rightarrow 2\text{Ni(OH)}_2(s) + 2 \text{OH}^- (aq) \]

**Net**

\[ \text{Cd}(s) + 2\text{NiO(OH)}(s) + 2 \text{H}_2\text{O(l)} \rightarrow \text{Cd(OH)}_2(s) + 2\text{Ni(OH)}_2(s) \]

Note: The number of reactions and the number of electrons hasn’t changed, but we’re *more completely* describing the physical and chemical form of the electrode components.
The cell contains a paste of NaOH – this provides the OH\(^{-}\) ions needed for the reaction, while also providing a medium to pass charge (electrolyte).

The anode consists of solid metal which is transformed into cadmium hydroxide.

The cathode consists of Ni\(^{3+}\) ions in a NiO(OH) paste which are transformed into nickel hydroxide.
It is because the products of the reaction are solids that the Ni-Cd battery can be recharged. The solid hydroxides are sticky, cling to the innards of the battery, and remain in place. If current is applied, the reaction can be driven backwards!
Batteries: The Nickel-Cadmium Battery

In a nickel-cadmium battery, we can recharge the battery by applying an electrical current from another source. The reaction is:

\[ \text{Cd(s) + 2NiO(OH)(s) + 2 H}_2\text{O(l) } \rightleftharpoons \text{Cd(OH)}_2\text{(s) + 2Ni(OH)}_2\text{(s)} \]

But most batteries we use aren’t rechargeable.

Why not?

What are the properties of some other typical batteries?
Billions upon billions of alkaline batteries are used each year. They are described by size and shape – AAA to D. Larger batteries have more “stuff”, and thus can run longer. But they all have the same voltage, because they’re all based on the same electrochemical cell.
Batteries: The Alkaline Battery

But they all have the same voltage, because they’re all based on the same electrochemical cell

Oxidation

\[
\text{Zn}(s) + 2 \text{OH}^- (aq) \rightarrow \text{Zn(OH)}_2(s) + 2 \text{e}^-
\]

Reduction

\[
2 \text{MnO}_2(s) + \text{H}_2\text{O}(l) + 2 \text{e}^- \rightarrow \text{Mn}_2\text{O}_3(s) + 2 \text{OH}^- (aq)
\]

Net

\[
\text{Zn}(s) + 2 \text{MnO}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{Zn(OH)}_2(s) + \text{Mn}_2\text{O}_3(s)
\]
But – the Mn$_2$O$_3$ is not sticky, and doesn’t remain attached to the electrode. This battery is not rechargeable.
Lithium-iodine batteries are particularly small and lightweight, but also very long-lived. Often used in pacemakers, where they can last for 10 years.

**Table 8.1 Some Common Galvanic Cells**

<table>
<thead>
<tr>
<th>Type</th>
<th>Voltage</th>
<th>Rechargeable?</th>
<th>Examples of Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alkaline</td>
<td>1.54</td>
<td>No</td>
<td>Flashlights, small appliances</td>
</tr>
<tr>
<td>Lithium–iodine</td>
<td>2.8</td>
<td>No</td>
<td>Camera batteries, pacemakers</td>
</tr>
<tr>
<td>Lithium ion</td>
<td>3.7</td>
<td>Yes</td>
<td>Laptop computers, cell phones, digital music players</td>
</tr>
<tr>
<td>Lead–acid (storage battery)</td>
<td>2.0</td>
<td>Yes</td>
<td>Automobiles</td>
</tr>
<tr>
<td>Nickel-cadmium (NiCd)</td>
<td>1.25</td>
<td>Yes</td>
<td>Consumer electronics</td>
</tr>
<tr>
<td>Nickel-metal hydride (NiMH)</td>
<td>1.25</td>
<td>Yes</td>
<td>Replacing NiCad for many uses; hybrid vehicles</td>
</tr>
<tr>
<td>Mercury</td>
<td>1.3</td>
<td>No</td>
<td>Formerly widely used in cameras, other appliances</td>
</tr>
</tbody>
</table>
Mercury batteries take advantage of the *high* density of Hg to be quite small: used in watches, hearing aids, calculators, etc.

Phased out in the 80s due to the toxicity of Hg.

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Batteries: The Lead-Acid Battery

Net:
\[
Pb(s) + PbO_2(s) + H_2SO_4(aq) \rightleftharpoons 2 PbSO_4(s) + 2 H_2O(l)
\]

The cathode is made of metallic lead, and the anode of lead dioxide.
The electrolyte is sulfuric acid.
This reaction, too, is reversible.
The lead sulfate product clings to the electrodes, so applied external voltage can reverse the reaction.
Batteries: The Lead-Acid Battery

Lead-acid batteries are referred to as “storage batteries”, because this charge-discharge cycle is so reliable.

These batteries were used in every automobile until quite recently.

The battery is discharged in order to start the engine.

Once the engine is running and burning gasoline, it turns an alternator which recharges the battery.

This process can continue for up to 5 years of normal driving.

After that time, enough of the lead sulfate product has been shaken off the plates that it can no longer recharge.
Batteries: The Lead-Acid Battery

Lead-acid batteries are also used in environments where vehicles cannot emit combustion products:
Indoor forklifts, golf carts, handicapped carts in airports, wheelchairs

However, lead is an environmental concern!
How do we dispose of the millions and millions of batteries which die each year?

There is a very successful recycling program in the U.S. – 97% of spent batteries are recycled

But environmentally healthier options are under investigation

A leading contender is the magnesium-acid battery