- 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

Electrochemistry: Some Definitions

In a nickel-cadmium battery, the reactions look something like this:

Oxidation $Cd \rightarrow Cd^{2+} + 2\ e^-$

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

Net $Cd + 2\ Ni^{3+} \rightarrow Cd^{2+} + 2\ Ni^{2+}$

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

Electrochemistry: Some Definitions

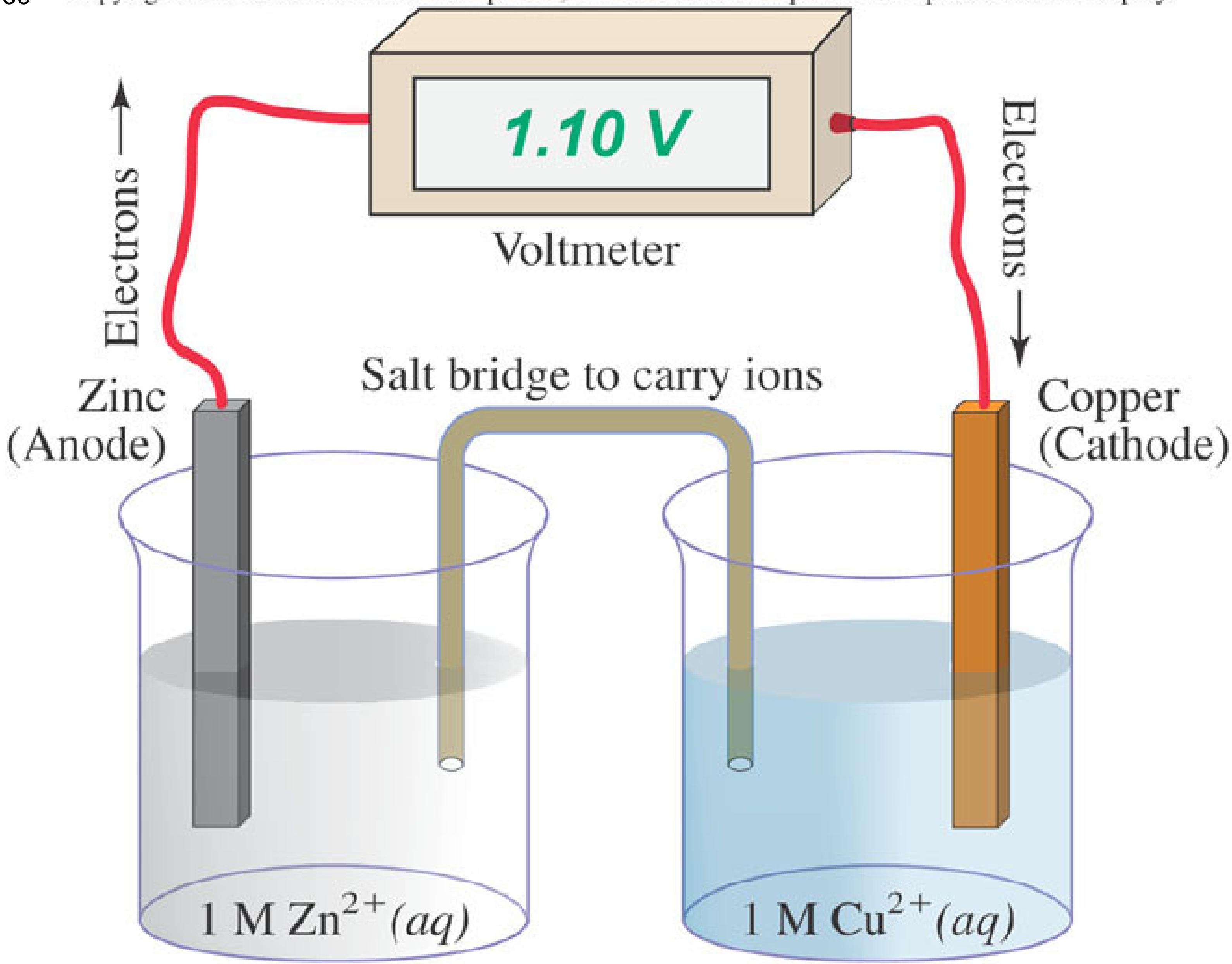
Electrodes are electrical conductors in the cell where chemical reactions take place

- 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)

Fig.08.p360



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

Batteries: The Nickel-Cadmium Battery

In a nickel-cadmium battery, the reactions **actually** look like this:

Oxidation

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

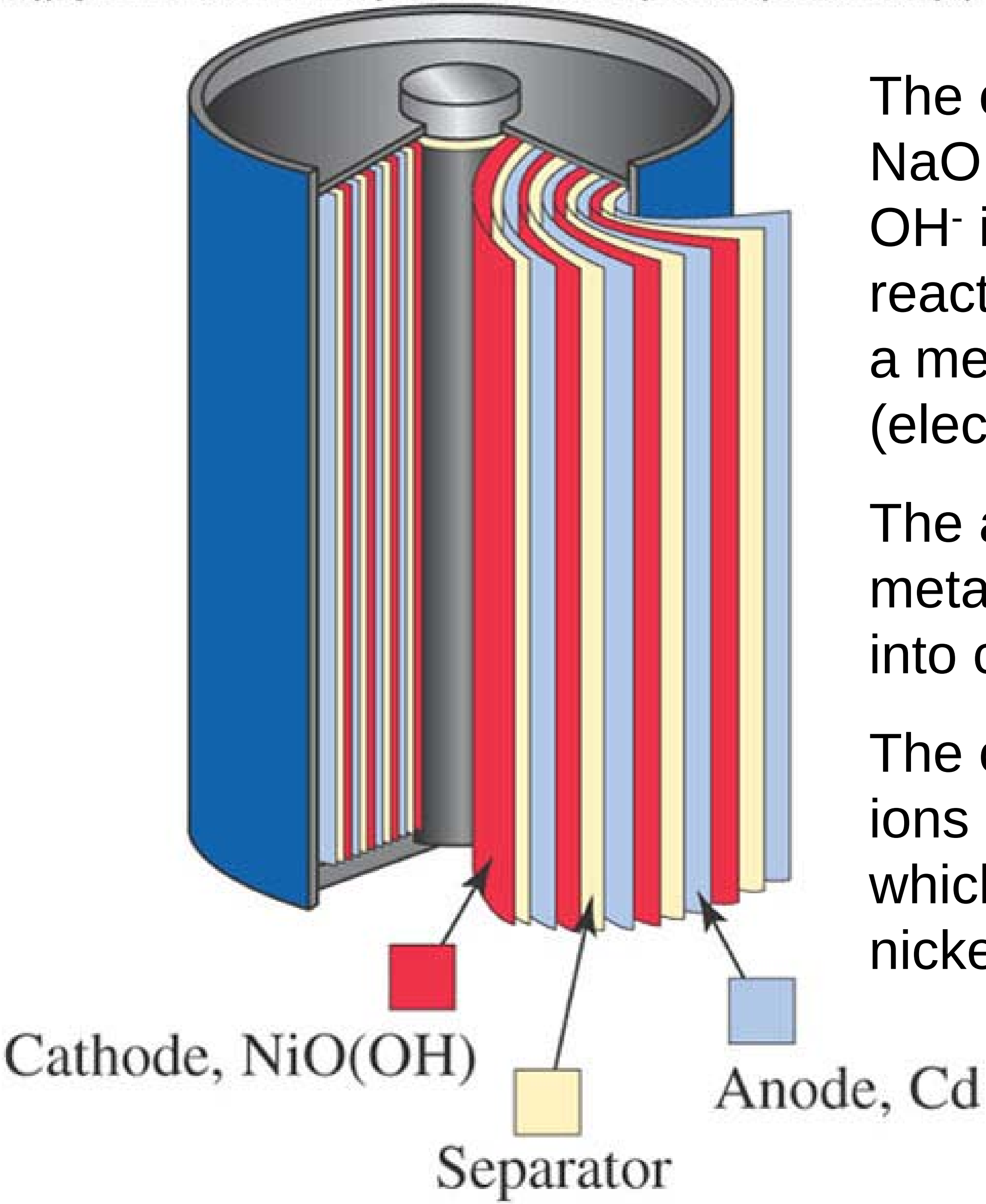
Reduction

$$2NiO(OH)(s) + 2\ H_2O(l) + 2\ e^{-} \rightarrow 2Ni(OH)_2(s) + 2\ OH\text{-}\ (aq)$$

Net

$$Cd(s) + 2NiO(OH)(s) + 2\ H_2O(l) \rightarrow Cd(OH)_2(s) + 2Ni(OH)_2(s)$$

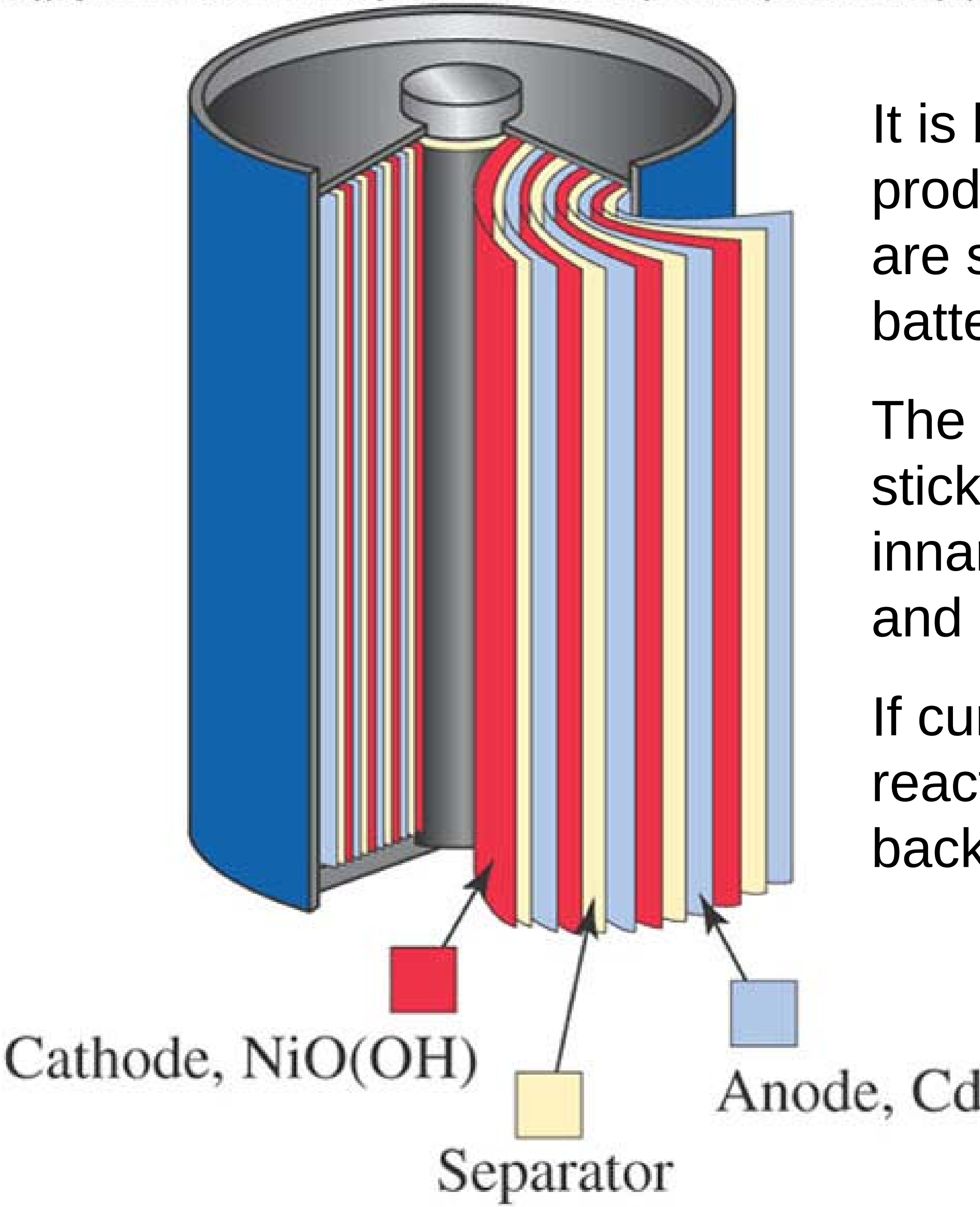
Note: The number of reactions and the number of electrons hasn’t changed, but we’re *more completely* describing the physical and chemical form of the electrode components



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

$Cd(s) + 2NiO(OH)(s) + 2\ H_2O(l) \rightleftarrows Cd(OH)_2(s) + 2Ni(OH)_2(s)$

But most batteries we use **aren't** rechargeable

Why not?

What are the properties of some other typical batteries?

Batteries: The Alkaline Battery

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

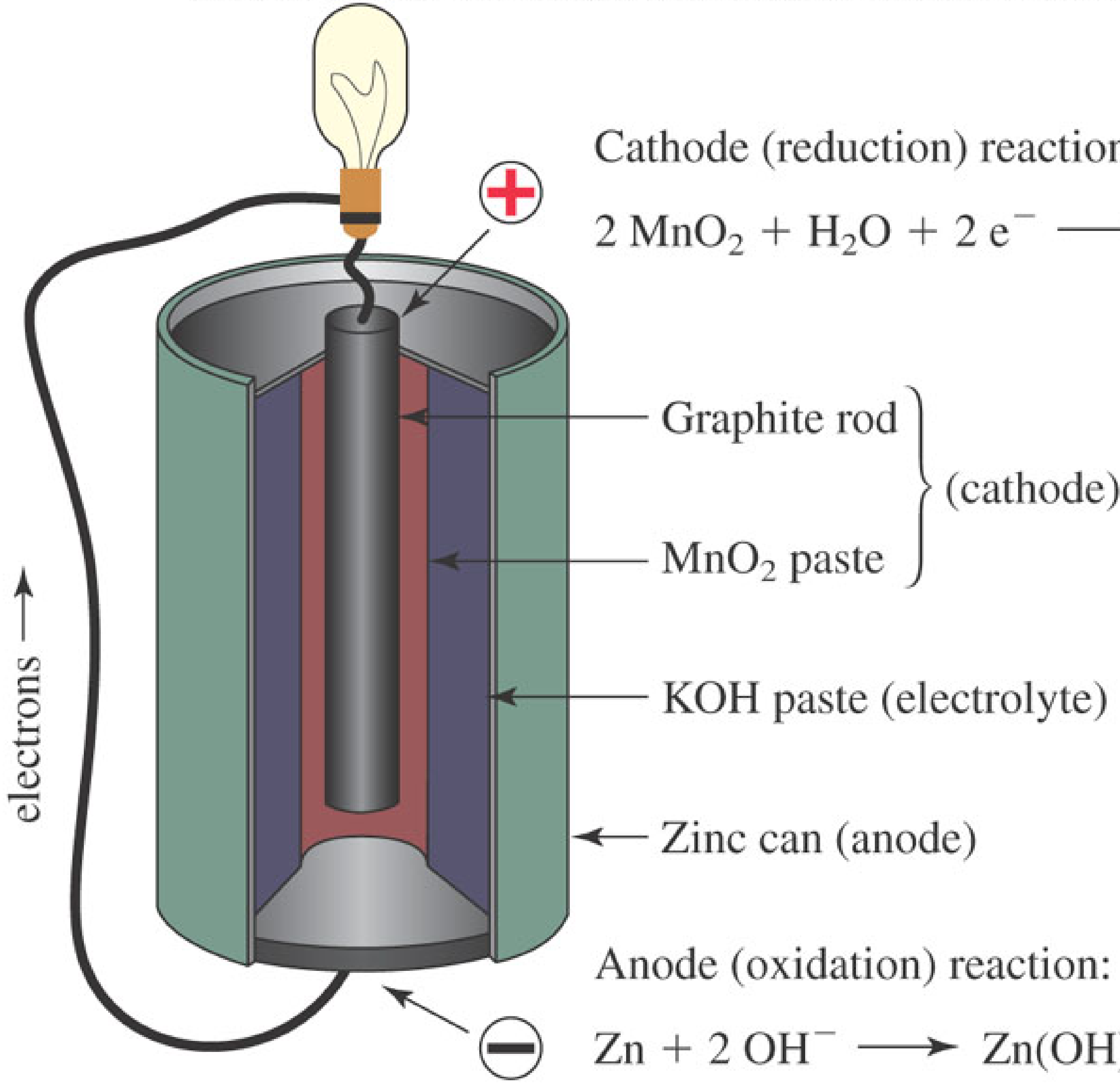
$$Zn(s) + 2\ OH^-(aq) \rightarrow Zn(OH)_2(s) + 2\ e^-$$

Reduction

$$2\ MnO_2(s) + H_2O(l) + 2\ e^- \rightarrow Mn_2O_3(s) + 2\ OH^-(aq)$$

Net

$$Zn(s) + 2\ MnO_2(s) + H_2O(l) \rightarrow Zn(OH)_2(s) + Mn_2O_3(s)$$



But – the Mn_2O_3 is not sticky, and doesn't remain attached to the electrode. This battery is not rechargeable

Table 8.1 Some Common Galvanic Cells

Type	Voltage	Rechargeable?	Examples of Uses
Alkaline	1.54	No	Flashlights, small appliances
Lithium–iodine	2.8	No	Camera batteries, pacemakers
Lithium ion	3.7	Yes	Laptop computers, cell phones, digital music players
Lead–acid (storage battery)	2.0	Yes	Automobiles
Nickel-cadmium (NiCd)	1.25	Yes	Consumer electronics
Nickel-metal hydride (NiMH)	1.25	Yes	Replacing NiCad for many uses; hybrid vehicles
Mercury	1.3	No	Formerly widely used in cameras, other appliances

Lithium-iodine batteries are particularly small and lightweight, but also very long-lived

Often used in pacemakers, where they can last for 10 years

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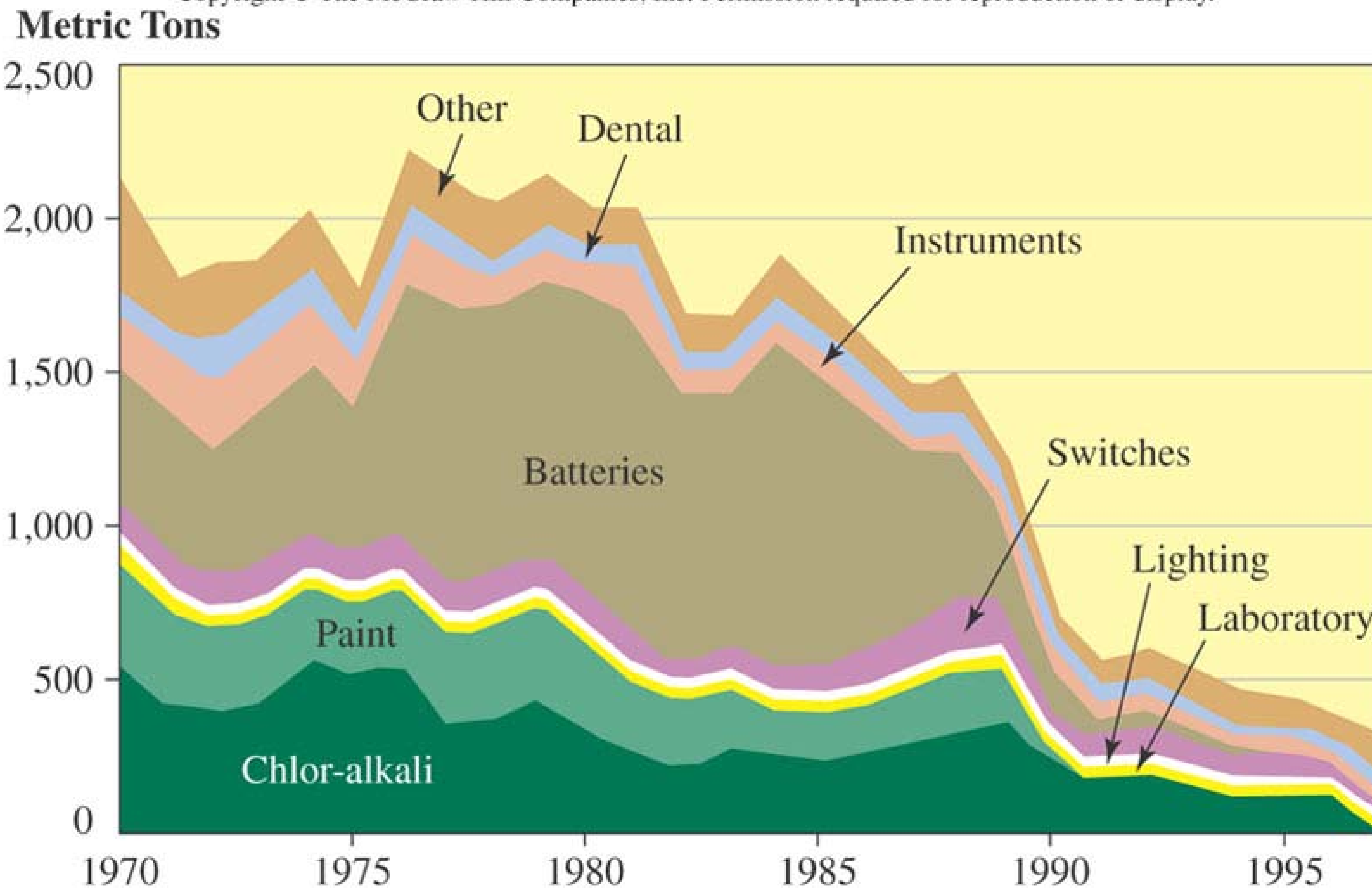
Mercury batteries take advantage of the *high* density of Hg to be quite small: used in watches, hearing aids, calculators, etc.

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

Fig.08.05


Metric Tons
2,500
2,000
1,500
1,000
500
0
1970
1975
1980
1985
1990
1995
Other
Dental
Instruments
Batteries
Switches
Lighting
Laboratory
Paint
Chlor-alkali

Batteries: The Lead-Acid Battery

Net:

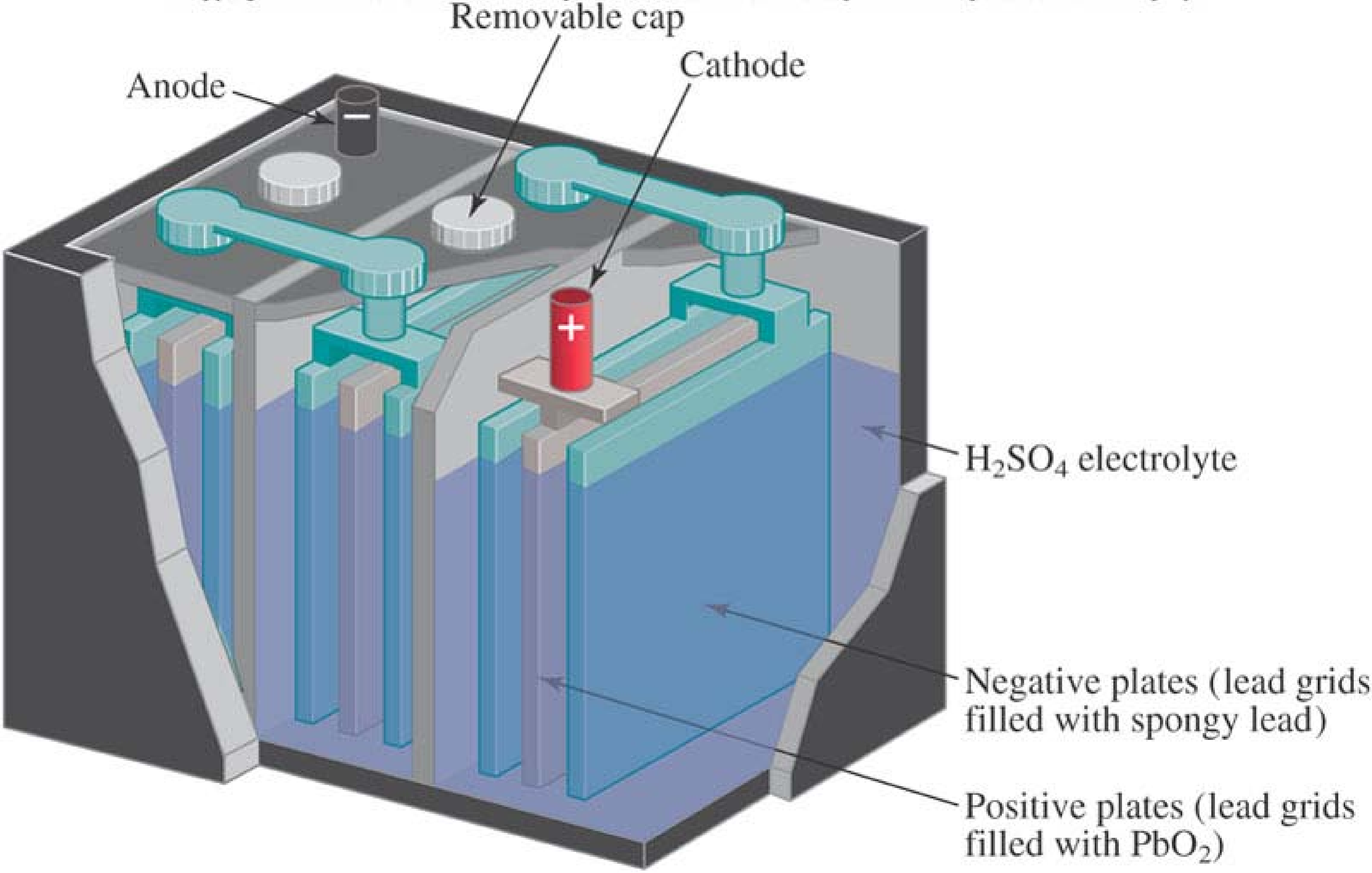
$Pb(s) + PbO_2(s) + H_2SO_4(aq) \rightleftharpoons 2\ PbSO_4(s) + 2\ H_2O(l)$

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


Removable cap
Cathode
Anode
H_2SO_4 electrolyte
Negative plates (lead grids filled with spongy lead)
Positive plates (lead grids filled with PbO_2)

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 succesful 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