EXPERIMENTER'S CORNER

By Forrest M. Mims

Do-It-Yourself Batteries

YOU can conduct a fascinating demonstration of electrical power generation by chemical means with a silver coin, a strip of magnesium and a piece of paper towel the size of a postage stamp. Dip the paper in lemon juice, place it over the coin and lay the magnesium strip on the paper. Then touch the cathode lead of a red LED to the magnesium. When the LED's anode lead is touched to the coin as in Fig. 1, the LED will glow brightly.

With the exception of the LED, this simple demonstration would have seemed fairly routine to pre-World War I experimenters. In those days many experimenters constructed their own primary and secondary (storage) cells. Commercial power cells were relatively expensive, and only a few kinds were available.

Today, literally hundreds of different kinds of batteries are available in a wide range of voltages and physical configurations. Nevertheless, for some special-purpose applications, a homemade battery may be more satisfactory than a commercial battery!

One example is powering a telemetry transmitter in an instrumented model rocket. The upward flight of such a rocket might last only a few seconds, yet a commercial battery could supply the necessary power continuously for days or even weeks. The penalty for this unnecessary capacity is excessive size and mass, both of which should be kept to a minimum.

This month, we will experiment with several electrochemical power cells that you can make from readily available materials. These cells are suitable for powering CMOS and other low-power circuits. Whether you assemble any of these cells or not, you may gain a better understanding of how conventional batteries work. If you do build cells of your own, you will certainly gain an appreciation for the convenience and drip-free operation of commercial power cells and batteries.

Some Definitions. Before proceeding any further, it is important to define a few basic terms:

Anode—The negative electrode of a cell.
Battery—Two or more electrically connected cells.
Cathode—The positive electrode of a cell.
Cell—A single two-electrode electrochemical generator.
Electrolyte—An ionized, electrically conductive paste, gel or liquid.
Primary Cell—A nonrechargeable cell.
Secondary Cell—A rechargeable cell.
Storage Cell—A secondary cell.

Electrochemical Generators. Alessandro Volta, an Italian physicist, invented the chemical generator. In March 1800, he demonstrated two of his generators before the Royal Society in London. One, called the Crown of Cups, consisted of a circular pattern of cups containing a solution of water and salt. One strip each of silver and zinc were immersed in each cup, and the zinc strip in one cup was connected to the silver strip in the adjacent cell. This arrangement formed a series connection of *wet cells*.

Volta's second generator was a stack of alternating disks of dissimilar metals separated by disks of paper soaked in brine. This device could produce more electromotive force in a smaller space than the clumsier Crown of Cups arrangement.

The electrochemical generators invented by Volta were used with little variation until about 1860, when other kinds of cells were developed. One such cell, patented by French scientist Georges LeClanche in 1868, was the predecessor of the modern zinc-carbon *dry cell*.

The basic design of the zinc-carbon dry cell (such as those used to power radios, flashlights and toys) has remained largely unchanged for more than sixty years. Each cell consists of a zinc cup or can (the anode) filled with a moist compound whose composition has changed through the years. One 1924 recipe called for a mixture of one ounce each of zinc-chloride and ammonium-chloride, two ounces of water, and three ounces of plaster of paris, which served as a filler. Sawdust was also used as a filler material in some early cells. A

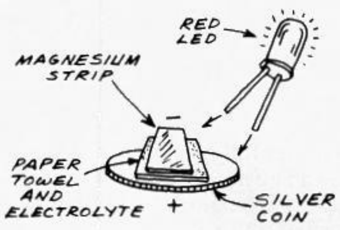

Fig. 1. An ultra-simple homemade power cell.

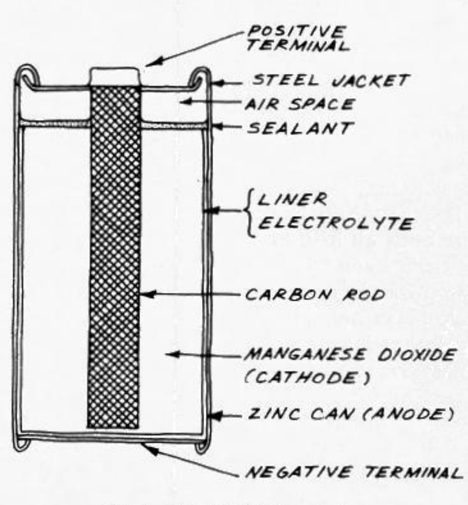

Fig. 2. Simplified internal view of a modern dry cell.

carbon rod inserted into the compound serves as the positive electrode.

The moist compound of the 1924 recipe served as the cell's electrolyte. In today's cells, the electrolyte is a paper liner impregnated with ammonium- or zinc-chloride that slides inside the zinc can. The space between the liner and the cell's carbon rod is packed with a mixture of granulated carbon and manganese dioxide. The latter compound serves as the cell's cathode. It is considered a *depolarizer* because it prevents *polarization*, the formation of an insulating layer of hydrogen bubbles around a cell's positive electrode.

Figure 2 is a pictorial view of the inner construction of a typical zinc-carbon dry cell. Most such cells are well sealed to prevent leakage which might occur should the zinc become corroded. The zinc seal also keeps the electrolyte from drying out. Drying of the electrolyte and subtle chemical reactions at the electrodes over time eventually degrade a cell whether or not it is used.

A Homemade Wet Cell. Figure 3 shows how you can make a simple wet cell from a plastic container such as a 35-mm film holder, a strip of copper and a strip of zinc. The

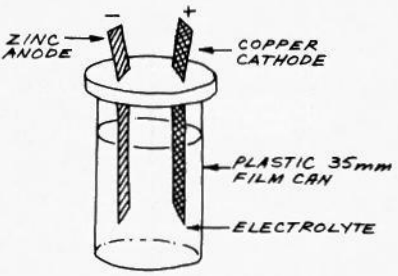

Fig. 3. Sketch of a homemade wet cell.

metal strips are inserted through slits cut in the container's cap or lid. The container is then filled with an electrolyte such as salt water or lemon juice, and the cap and the electrodes are installed. This cell will produce about 0.7 volt.

If you connect a voltmeter to the cell's electrodes and pull the electrodes partially out of the electrolyte, the output voltage will remain unchanged. Even when only a few millimeters of each electrode remain immersed in the electrolyte, the output voltage will remain unchanged. The cell's capacity to deliver current, however, is directly proportional to the area of the electrodes immersed in the electrolyte.

Electrode Materials. Any two dissimilar metals immersed in a suitable electrolyte will generate a voltage. Here are the voltages I measured for all possible pairs selected from the following group: a copper penny, a nickel, a silver dime, a magnesium strip, a zinc strip, and aluminum foil.

Cathode (+)	Anode (−)	Open-Circuit Voltage
Nickel	Copper	0.04
Magnesium	Zinc	0.05
Aluminum	Zinc	0.15
Silver	Nickel	0.19
Silver	Copper	0.20
Copper	Aluminum	0.70
Nickel	Aluminum	0.70
Copper	Zinc	0.72
Aluminum	Magnesium	0.78
Nickel	Zinc	0.81
Silver	Aluminum	0.84
Silver	Zinc	1.01
Nickel	Magnesium	1.44
Copper	Magnesium	1.45
Silver	Magnesium	1.65

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For these measurements, I used as a combined electrolyte and separator several layers of paper towel soaked in salt water. The values you measure may differ slightly from mine, particularly if you use an acid electrolyte, in which case the values will be higher.

Note that the highest voltages are produced by pairing magnesium with nickel, copper or silver. Nickel and copper can be found in pocket change, but silver coins have not been minted in the U.S. since the mid-Sixties and are rarely found in everyday circulation. You can purchase magnesium strips at toy and hobby shops that sell Perfect brand chemicals.

Magnesium is highly reactive. I tried a magnesium strip in a lemon-juice wet cell and found that although the cell could easily power a LED, the magnesium was soon covered by a frothy layer of hydrogen bubbles. The cell functioned well in spite of the bubbles until the reaction with the citric acid in the lemon juice coated the magnesium with a black film.

Zinc is the next-best substitute for magnesium. You can get free zinc by cutting open a discarded zinc-carbon flashlight cell. If the cell is covered by an outer steel jacket, use pliers or diagonal cutters to peel it off. Be careful! The edges of the metal envelope will be very sharp.

When the zinc can has been exposed (it may be covered with a layer of paper or black pitch), secure the cell in a vise and use a hacksaw to remove the top half-inch of the cell. Remove the carbon rod, the filler compound and the electrolyte-impregnated paper liner from inside the can. The carbon in the compound will stain clothing, so be careful. Watch out for the sharp edges of the zinc can and clean any remaining compound from the can with detergent, water and an old toothbrush.

When the zinc is clean, use a file to remove the sharp edges

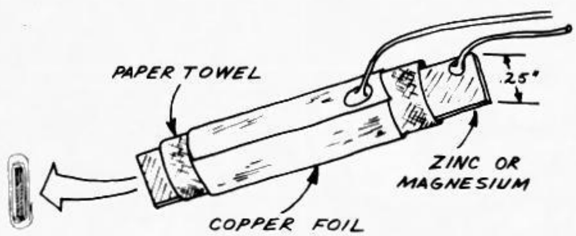

Fig. 4. How to make a simple 1.45-volt "moist" cell.

left by the hacksaw. Then cut the can into strips with shears or a nibbling tool. Remove any corrosion from the strips with sandpaper.

Homemade "Moist" Cells. Figure 4 shows a simple 1.45-volt cell made from a 1/4-inch wide strip of magnesium or zinc wrapped with two layers of paper towel previously dipped in a solution of salt and water. A piece of copper foil (available at craft and hobby shops) the size of a postage stamp is wrapped over the paper towel.

For best results, the paper towel should be dried before the cell is assembled. When the cell is to be used, it can be activated by dipping it in water. Alternatively, a few drops of water can be applied to the exposed ends of the paper towel.

The cell shown in Fig. 4 is merely one of many possible configurations. You can make round, square or triangular cells. You can even cut the zinc or magnesium anode into long, nar-

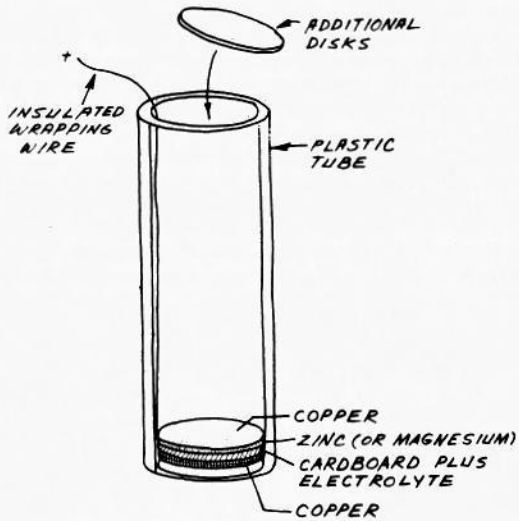

Fig. 5. Construction of a multi-cell stacked battery using zinc or magnesium.

row strips and make ultra-thin, cylindrical cells. You can increase the current capacity of a cell by increasing the area of its electrodes. Two or more cells can be connected in series to achieve higher voltages.

If a discharged cell is disconnected, in time it will gradually recover. Add moisture, and the cell will again deliver power. After several discharge cycles, you can rejuvenate a cell by unwrapping the copper foil and cleaning both the anode and cathode with steel wool. Reassemble the cell with a fresh, salt-impregnated separator layer.

There are two ways to attach wires to the cell. The simplest is to use miniature clip leads. I prefer to solder short lengths of wrapping wire to the electrodes prior to assembly. Copper foil is easily soldered. Zinc must be sanded for best results. Solder will not adhere to magnesium, so you will have to use a clip lead if you use this anode material.

Homemade Stacked Batteries. You can assemble a miniature version of Volta's stacked battery, which was called a *Voltaic pile*, with the help of a 1/4-inch paper punch. Punch a dozen or so holes in a piece of *thin* cardboard like that used for shoe boxes and soak the cardboard disks in salt water or lemon juice. Then punch an identical number of disks out of sheets of copper foil and magnesium or zinc. Solder a six-inch length of wrapping wire to one copper disk and one zinc disk. If you use magnesium, make an extra copper disk and solder a wire to it.

For best results, assemble the cell inside a hollow plastic tube. I used a small tube which originally contained the point of a drafting pen. Flexible tubing can also be used.

Install the copper disk with an attached wire lead first. Then, install alternating disks of cardboard, zinc (or magnesium), copper, cardboard, etc. The final disk should be zinc with an attached wire. If magnesium is used in place of zinc, top off the stack with the other copper disk to which a lead has been attached.

You may need to press the disks lightly against the end of the tube to achieve maximum output from the battery. Too much pressure, however, will squeeze electrolyte from the cardboard disks. Free electrolyte can short adjacent cells and reduce the battery's output voltage.

Incidentally, you will find it very helpful to use pointed tweezers when assembling a battery like this. Also, be sure to blot excess electrolyte from the cardboard disks before installing them in the tube.

Figure 5 shows how a 12-cell stack is assembled. This battery delivers 5.8 volts open circuit and is able to drive a LED with built-in flasher. The load of the LED and flasher circuit drops the voltage from the battery to a few volts, so the LED is not very bright. I used zinc and copper disks and lemon juice electrolyte.

Unfortunately, the battery in Fig. 5 is not suitable for post-assembly water activation. Adding water to the battery would short all the cells together.

Additional Reading. Because of the volume of mail I receive, it will not be possible to answer individual reader's questions about homemade batteries. Fortunately, however, there are many good books on the subject.

For best results, visit a library which has lots of old books. Some large libraries keep such books in a separate section. Small-town libraries tend to keep such books longer than big libraries.

One typical example is *Electricity and Magnetism and Their Applications* by Dugald Jackson and John Jackson. This text was originally published by the Macmillan Company in 1902. It was revised by N. Henry Black in 1919 and republished in 1920. It contains an excellent chapter on early battery technology.

Another old book is *The Amateur Electrician's Handbook* by A. Frederick Collins (Thomas Y. Crowell Company, 1924). It contains instructions for making everything from wet cells to X-ray systems. I would not advise you to try the latter because much has been learned about the hazards of ionizing radiation since 1924.

Many newer books also discuss batteries. One is *Basic Electronics* by Abraham Marcus and Samuel Gendler (Prentice-Hall, 1971). Another is *Handbook for Electronics Engineering Technicians* by Milton Kaufman and Arthur Seid-