In every electrochemical process, whether spontaneous or not, a certain amount of electric charge is transferred during the oxidation and reduction. The half-reactions we have written for electrode processes include the electrons which carry that charge. It is possible to measure the rate at which the charge is transferred with a device called an ammeter.

An ammeter measures the current flowing through a circuit. The units of current are amperes (A) (amps, for short). Unlike a voltmeter, ammeters allow electrons to pass and essentially "clock" them as they go by. The amount of electric charge which has passed through the circuit can then be calculated by a simple relationship:

\[
\text{Charge} = \text{current} \times \text{time} \quad \text{OR} \quad \text{Coulombs} = \text{amps} \times \text{seconds}
\]

This enables us to connect reaction stoichiometry to electrical measurements. The principles underlying these relationships were worked out in the first half of the 19th century by the English scientist, Michael Faraday.

The diagram shows how voltage and current might be measured for a typical galvanic cell but the arrangement is the same for any electrochemical cell. Notice that the voltmeter is placed across the electron conduit (i.e., the wire) while the ammeter is part of that conduit. A good quality voltmeter can be used in this way even though it might appear to be "shorting out" the circuit. Since electrons cannot pass through the voltmeter, they simply continue along the wire.

Both the voltmeter and ammeter are polarized. They have negative and positive terminals marked on them. Electrons are "expected" only in one direction. This is important in measurements of direct current (DC) such as comes out of (or goes into) electrochemical cells.

**Faraday's law of electrolysis** might be stated this way: the amount of substance produced at each electrode is directly proportional to the quantity of charge flowing through the cell. Of course, this is somewhat of a simplification. Substances with different oxidation/reduction changes in terms of the electrons/atom or ion will not be produced in the same molar amounts. But when those additional ratios are factored in, the law is correct in all cases.
Example 1

The electrolysis of molten MgCl$_2$ is the last step in the isolation of magnesium metal from seawater. How many amps are needed to produce 35.6 g of Mg in 2.50 hours?

Example 2

A galvanic cell consisting of standard concentrations: Cu|Cu$^{2+}$|Ag$^+$|Ag, is connected to a small light bulb for 40.0 minutes. During that time the average current drawn by the bulb is 0.12 A. How many grams of copper dissolve from the anode during this time?

Example 3

A lead storage battery like one used to start your car contains lead and lead(IV) oxide electrodes in a sulfuric acid electrolyte. The electrode reactions are:

Anode: $\text{Pb} + \text{SO}_4^{2-} \rightarrow \text{PbSO}_4 + 2 \text{e}^-$

Cathode: $\text{PbO}_2 + 4 \text{H}^+ + \text{SO}_4^{2-} + 2 \text{e}^- \rightarrow \text{PbSO}_4 + 2 \text{H}_2\text{O}$

A typical automobile battery might be rated at "100 ampere-hours". This means it can deliver a current of 1.0 A for 100 hours.

How many grams of Pb are oxidized for the full rating?

Example 4

How many minutes does it take to produce 10.0 L of oxygen gas (measured at 755 mmHg and 25.0°C) by electrolyzing neutral water with a current of 1.3 A? What mass of hydrogen forms at the same time?

Example 5

The violet vanadium(II) ion, V$^{2+}$, can be produced by electrolysis of a solution containing the green V$^{3+}$ ion. How long will it take to accomplish this in 0.35 L of a 0.50 M V$^{3+}$ solution with a current of 0.26 A?

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