Copper bottomed

Topics

Metals, reactivity series

In the 18th century, it was common for wooden ships to have their hulls sheathed with copper as a protection against barnacles and other marine organisms which could slow the ship and attack the wood below the waterline. This is where the expression ‘copper-bottomed’ meaning ‘reliable’ originates. However, there was a problem with this technique. Iron bolts were used to hold the various parts of the wooden hulls together and if these came into contact with the copper of the sheathing immersed in sea water, the ideal conditions for bimetallic corrosion were in place. Here the more reactive metal, iron, corrodes faster than normal when in contact with a less reactive metal. This led to rapid deterioration of the hull bolts and let to the possibility that sections of the hull could come away with disastrous consequences. There are stories (uncorroborated) of ships at sail whose copper bottom fell off, leaving the ship top heavy and causing it to capsize.

It is unlikely that this phenomenon was ever solely responsible for the loss of a ship but there is one well-documented instance where it might have been a contributory factor. Maintenance work was being carried out on the warship the Royal George in 1782. The captain caused her to heel over to facilitate this by moving all the guns over to one side. Shortly after this the ship sank with little warning and some reports indicate that a loud crack was heard at the time. The cause of the sinking was later established to be that part of the ship’s frame collapsed and it is suggested that this might have been caused by corroded bolts.

Another example of bimetallic corrosion in sea water is that of sacrificial protection of steel hulls of ships (and also oil rigs and pipelines). Here blocks of zinc (or sometimes magnesium) are attached to the steel (*ie* mostly iron) hull of the ship. Here the more reactive zinc corrodes away faster and protects the iron.

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

proceeds to the right and the electrons that are released force the equilibrium below to the left

$$Fe(s) + aq \rightleftharpoons Fe^{2+}(aq) + 2e^-$$

The zinc corrodes away and has to be regularly replaced.