C19 Balancing the Redox Reaction

Rob Lederer Example that was at the top of the whiteboard is a redox reaction

$$\ldots Fe + \ldots HCl \rightarrow \ldots HFeCl_4 + \ldots H_2$$

This reaction has a complication that you would not see on the AP Chem exam.
For those of you who want to know how to balance it anyway here's how it would be balanced.

First find the redox numbers in the unbalanced reaction:

$$\overset{0}{Fe} + \overset{+1}{H}\overset{-1}{Cl} \rightarrow \underset{+1}{H}\underset{+3}{Fe}\underset{-4}{Cl_4} + \overset{0}{H_2}$$

The redox half reactions:

$2\,e^- + 2\,H^+ \rightarrow H_2$ Reduction
$Fe^0 \rightarrow Fe^{+3} + 3\,e^-$ Oxidation

There is a complication in that there are some H atoms are not reduced. We will include those later.

Balance the two half reactions so that the electrons reduced equal the electrons oxidized.

$6\,e^- + 6\,H^+ \rightarrow 3\,H_2$ Reduction (multiplied by 3 to get 6 e^- reduced)
$2\,Fe^{+3} \rightarrow 2\,Fe^{+3} + 6\,e^-$ Oxidation (multiplied by 2 to get 6 e^- oxidized)

Therefore, in the balanced reaction, there must be 6 H for every 2 Fe
So, these coefficients would be placed in the balanced reaction taking into account only the H atoms that are oxidized.

$$2\,Fe + 6\,HCl \rightarrow 2\,HFeCl_4 + 3\,H_2$$

The redox part is balanced but because there are 2 H atoms that are not reduced, we need to put them in completely balance the reaction. The two HCl atoms that are not part of the redox reaction can be added to the reaction to correctly balance it.

$$2\,Fe + 6\,HCl + 2\,HCl \rightarrow 2\,HFeCl_4 + 3\,H_2$$

$$2\,Fe + 8\,HCl \rightarrow 2\,HFeCl_4 + 3\,H_2$$