In the **Standard Reduction Table**, the voltages for half cell reactions are given. The half-cell reaction : \(2H^+ + 2e^- \rightarrow H_2\) is assigned a voltage of 0.0 volts. This is arbitrary as the half-cell reaction : \(Cu^{2+} + 2e^- \rightarrow Cu\) which is assigned a voltage of 0.34 volts could have been assigned a voltage of 0.0 volts. This would make the hydrogen half-cell voltage equal to -0.34 volts. These half-cell voltages are **relative**.

These voltages do not mean anything by themselves, as a reduction cannot occur without an oxidation. The important thing is the difference in voltages between the reduction half-reaction and the oxidation half-reaction. We can then find the voltage that a particular electrochemical cell can produce under standard conditions.

**Examples :**

1) For a silver-nickel cell we have the following voltages :

\[
\begin{align*}
Ag^+ + e^- & \rightarrow Ag & +0.80 \text{ volts reduction} \\
Ni & \rightarrow Ni^{2+} + 2e^- & +0.26 \text{ volts oxidation}
\end{align*}
\]

The voltages are found in the table of reduction potentials. To find the voltage of the oxidation half-cell reaction, simply change the sign so that the nickel half-reaction has a voltage of +0.26 instead of -0.26 volts. The voltage of the cell can be found by simply **adding** the two voltages. +0.80 +0.26 = +1.06 volts under standard conditions. The + sign means that the reaction is **spontaneous**. The standard conditions are : the cell electrodes are immersed in solutions with a concentration of 1.0 M, the temperature is 25°C and the pressure is 101.3 kPa (1.0 atmospheres).

2) In the above reaction : \(2Ag^+ + Ni \rightarrow Ni^{2+} + 2Ag\), the overall voltage is positive so the reaction is **spontaneous**, therefore, it will proceed. For the reaction : \(Ni + Fe^{2+} \rightarrow Ni^{2+} + Fe\), the cell voltage is -0.19 volts. When the cell voltage is negative, the reaction will be **non-spontaneous**. In this case energy must be supplied to the cell to make the reaction proceed.
Problems: 1) Using the Standard Reduction Potential tables, for the following redox reactions, find the cell potential, and indicate whether the reactions are spontaneous, non-spontaneous, or not possible.

a) $\text{Ni}^{2+} + \text{Fe} \rightarrow \text{Ni} + \text{Fe}^{2+}$

b) $\text{Ag} + \text{H}^+ \rightarrow \text{Ag}^+ + \text{H}_2$

c) $\text{Co} + \text{Fe}^{3+} \rightarrow \text{Co}^{2+} + \text{Fe}^{2+}$

d) $\text{I}^- + \text{MnO}_4^- \rightarrow \text{I}_2 + \text{Mn}^{2+}$

e) $\text{Ag} + \text{Cl}^- \rightarrow \text{Ag}^+ + \text{Cl}_2$

f) $\text{I}_2 + \text{Cl}^- \rightarrow \text{I}^- + \text{Cl}_2$

g) $\text{H}_2 + \text{S}_2\text{O}_8^{2-} \rightarrow \text{H}^+ + \text{SO}_4^{2-}$

Answers: 1)a) +0.19V, spontaneous, b) -0.80V, non-spontaneous, c) +1.05V, spontaneous, d) +0.97V, spontaneous, e) not possible (it is not a redox reaction), f) -0.82V, non-spontaneous, g) +2.01V, spontaneous.