Reduction of hydronium to form H_2

At what value pH will Ni be capable of reducing H^+ to H_2?

$$Ni^{2+} + 2\,e^- \rightarrow Ni \qquad\qquad E^o_{red} = -0.250\,V$$

Reduction of hydronium to form H_2

At what value pH will Ni be capable of reducing H^+ to H_2?

$$Ni^{2+} + 2\,e^{-} \rightarrow Ni \qquad E^{o}_{red} = -0.250\,V$$

Solution: Under standard conditions the redox reaction and cell potential are:

$$Ni + 2\,H^{+} \rightarrow Ni^{2+} + \; H_2 \qquad E^{o}_{cell} = +0.25$$

We can use this equation to write the Nernst equation:

$$E_{cell} = E^{o}_{cell} - \frac{RT}{nF} ln \frac{[Ni^{+}]P_{H_2}}{[H^{+}]^{2}}$$

Here we assume that Ni^+ and the H_2 pressure are under standard conditions of 1 M and 1 bar, respectively. We can also make the following substitution to convert the logarithm to base 10:

$$2.303\; log_{10}[H^{+}] = ln[H^{+}]$$

Reduction of hydronium to form H_2

At what value pH will Ni be capable of reducing H^+ to H_2?

$$Ni^{2+} + 2\ e^- \rightarrow Ni \qquad\qquad E^o_{red} = -0.250\ V$$

Consequently we can write the Nernst equation as:

$$E_{cell} = 0.25 + 0.0257(2.303)log_{10}[H^+]$$

Finally, we can use the definition of pH to write:

$$E_{cell} = 0.25 - 0.0591\ pH$$

To find the pH at which the reaction will proceed we solve for the pH when $E_{cell} = 0$.

$$pH = \frac{0.25}{0.0591} = 4.23$$