

# Reduction of hydronium to form H<sub>2</sub>

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



We can use this equation to write the Nernst equation:

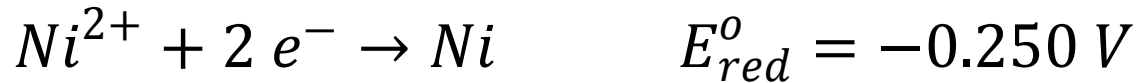
$$E_{cell} = E_{cell}^{\circ} - \frac{RT}{nF} \ln \frac{[Ni^{2+}]P_{H_2}}{[H^{+}]^2}$$

Here we assume that Ni<sup>2+</sup> and the H<sub>2</sub> 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:

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

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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:

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

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

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