Question

Represent the cell in which the following reaction takes place

$$Mg(s) + 2Ag^{+}(0.0001M) \longrightarrow Mg^{2+}(0.130M) + 2Ag(s)$$

Calculate its E_{cell} if $E^{\circ}_{cell} = 3.17V$

Answer

$$Mg(s) \mid Mg^{2+}(aq)(0.130M) \parallel Ag^{+}(aq)(0.0001M) \mid Ag(s)$$

Nernst equation (at 298 k)

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \frac{[Mg^{2+}]}{[Ag^{+}]^2}$$

$$= 3.17 - \frac{0.059}{2} \log \frac{0.130}{(0.0001)^2}$$

$$= 3.17 - \frac{0.059}{2} \log 13 \times 10^6$$

$$= 3.17 - 0.21$$

$$= 2.96V$$

$\log 13 \times 10^6$

$\log 13 + \log 10^6$

$\log 2 \times 2 \times 3 + 6 \log 10$

$\log 2 + \log 2 + \log 3 + 6 \log 10$

$0.3 + 0.3 + 0.48 + 6(1)$

$\log 2 = 0.3$ $\log 5 = 0.7$

$\log 3 = 0.48$ $\log 7 = 0.85$

Question

Calculate the emf of the cell in which the following reaction takes place

$$Ni(s) + 2Ag^{+}(0.002M) \longrightarrow Ni^{2+}(0.160M) + 2Ag(s)$$

Given: $E^{\circ}_{cell} = 1.05V$

Answer

By Nernst equation (at 298 k)

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \frac{[Ni^{2+}]}{[Ag^{+}]^2}$$

$$= 1.05 - \frac{0.059}{2} \log \frac{0.160}{}$$

$\log \frac{0.16}{}$

$$= 1.05 - \frac{0.059}{2} \log \frac{0.160}{(0.002)^2}$$

$$= 1.05 - 0.1357$$

$$= 0.914V$$

$$\log \frac{0.16}{0.002 \times 0.002}$$

$$\log \frac{16 \times 10^6}{100 \times 2 \times 2}$$

$$\log 4 \times 10^4$$

$$\log 2 \times 2 + \log 10^4$$

$$\log 2 + \log 2 + 4 \log 10$$

$$= 0.3 + 0.3 + 1$$

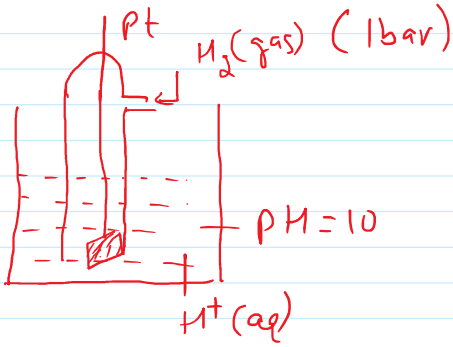
$$\log ab = \log a + \log b$$

$$\log \frac{a}{b} = \log a - \log b$$

$$\log a^b = b \log a$$

Question

Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.



Answer

Electrode reaction

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

Apply Nernst equation (at 298 k)

$$E_{H^+/H_2} = E^\circ_{H^+/H_2} - \frac{0.059}{n} \log \frac{P_{H_2}}{[H^+]^2}$$

$$= 0 - \frac{0.059}{2} \log \frac{1}{[H^+]^2}$$

$$= -\frac{0.059}{2} \left(\log [H^+]^2 \right)$$

$$= -\frac{0.059}{2} (-2) \log [H^+] \qquad (\log a^b = b \log a)$$

$$= -0.059 \left(-\log [H^+] \right)$$

$$\boxed{E_{H^+/H_2} = -0.059 pH} \qquad (\text{at } 298 k,\ P_{H_2} = 1 bar)$$

$$= -0.059 \times 10$$

$$= -0.59V$$

Question

Write the Nernst equation and emf of the following cells at 298 k:

i) $Mg(s) | Mg^{2+} (0.001M) || Cu^{2+} (0.0001M) | Cu(s)$

$E^{o}_{Mg^{2+}/Mg} = -2.36V$ $\qquad$ $E^{o}_{Cu^{2+}/Cu} = 0.34V$

ii) $Fe(s) | Fe^{2+} (0.001M) || H^{+} (1M) | H_2(g)(1bar) | Pt(s)$

$E^{o}_{Fe^{2+}/Fe} = -0.44V$

iii) $Sn(s) | Sn^{2+} (0.050M) || H^{+} (0.020M) | H_2(g)(1bar) | Pt(s)$

$E^{o}_{Sn^{2+}/Sn} = -0.14V$

iv) $Pt(s) | Br^{-}(0.010M) | Br_2(l) || H^{+} (0.030M) | H_2(g) (1bar) | Pt(s)$

$E^{o}_{Br_2/Br^{-}} = 1.09V$

Answer:

i) Anode half cell reaction

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^{-}$ $\qquad$ $E^{o}_{Mg^{2+}/Mg} = -2.36V$

Cathode half cell reaction

$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$ $\qquad$ $E^{o}_{Cu^{2+}/Cu} = 0.34V$

Cell reaction

$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

$E^{o}_{cell} = E^{o}_{Cu^{2+}/Cu} - E^{o}_{Mg^{2+}/Mg}$

$= 0.34 - (-2.36) = 2.7V$

By Nernst equation (at 298k)

$E_{cell} = E^{o}_{cell} - \frac{0.059}{n} \log \frac{[Mg^{2+}]}{[Cu^{2+}]}$

$= 2.7 - \frac{0.059}{2} \log \frac{0.001}{0.0001}$

$= 2.67V$

ii) Anode half cell reaction

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ $\qquad E^\circ_{Fe^{2+}/Fe} = -0.44V$

Cathode half cell reaction

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $\qquad E^\circ_{H^+/H_2} = 0V$

Cell reaction

$Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$

$E^\circ_{cell} = E^\circ_{H^+/H_2} - E^\circ_{Fe^{2+}/Fe}$

$= 0 - (-0.44) = 0.44V$

By Nernst equation (at 298K)

$E_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log \frac{[Fe^{2+}]}{[H^+]^2} P_{H_2}$

$= 0.44 - \frac{0.059}{2} \log \frac{0.001 \times 1}{1^2}$

$= 0.53V$ (approx)

iii) Anode half cell reaction:

$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$ $\qquad E^\circ_{Sn^{2+}/Sn} = -0.14V$

Cathode half cell reaction:

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $\qquad E^\circ_{H^+/H_2} = 0V$

Cell reaction

$Sn(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$

$E^\circ_{cell} = E^\circ_{H^+/H_2} - E^\circ_{Sn^{2+}/Sn}$

$= 0 - (-0.14) = 0.14V$

By Nernst equation

$E_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log \frac{[Sn^{2+}] P_{H_2}}{[H^+]^2}$

$= 0.14 - \frac{0.059}{2} \log \frac{0.05 \times 1}{(0.02)^2}$

$= 0.08V$ (approx)

2) Anode half cell reaction

$2Br^-(aq) \rightarrow Br_2(l) + 2e^-$ $\qquad E^\circ_{Br_2/Br^-} = 1.09V$

Cathode half cell reaction

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $\qquad E^\circ_{H^+/H_2} = 0V$

Cell reaction:

$2Br^-(aq) + 2H^+(aq) \rightarrow Br_2(l) + H_2(g)$

$E^\circ_{cell} = E^\circ_{H^+/H_2} - E^\circ_{Br_2/Br^-}$

$= 0 - 1.09 = -1.09V$

By Nernst equation (at 298K)

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log \frac{P_{H_2}}{[Br^-]^2 [H^+]^2}$$

$$= -1.09 - \frac{0.059}{2} \log \frac{1}{(0.01)^2 (0.03)^2}$$

$= -1.3V$