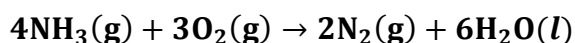


1. (10%) Consider the oxidation of ammonia:

(a) Calculate the ΔG° for the reaction.

(b) If this reaction were used in a fuel cell, what would be the standard cell potential?

Sol.

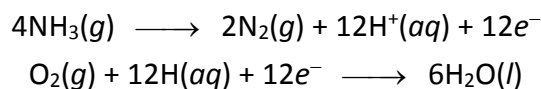
(a) We can calculate ΔG° from standard free energies of formation.

$$\Delta G^\circ = 2\Delta G_f^\circ(\text{N}_2) + 6\Delta G_f^\circ(\text{H}_2\text{O}) - [4\Delta G_f^\circ(\text{NH}_3) + 3\Delta G_f^\circ(\text{O}_2)]$$

$$\Delta G = 0 + (6)(-237.2 \text{ kJ mol}^{-1}) - [(4)(-16.6 \text{ kJ mol}^{-1}) + 0]$$

$$\Delta G = -1356.8 \text{ kJ mol}^{-1}$$

(b) The half-reactions are:

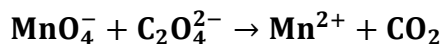


The overall reaction is a 12-electron process. We can calculate the standard cell emf from the standard free energy change, ΔG° .

$$\begin{aligned} \Delta G^\circ &= -nFE_{\text{cell}}^\circ \\ E_{\text{cell}}^\circ &= \frac{-\Delta G^\circ}{nF} = \frac{-\left(\frac{-1356.8 \text{ kJ}}{1 \text{ mol}} \times \frac{1000 \text{ J}}{1 \text{ kJ}}\right)}{(12)(96500 \text{ J V}^{-1} \text{ mol}^{-1})} = 1.17 \text{ V} \end{aligned}$$

2. (10%) Oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$) is present in many plants and vegetables.

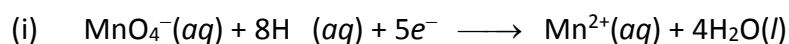
(a) Balance the following equation in acid solution:

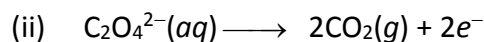


(b) If a 1.00 g sample of $\text{H}_2\text{C}_2\text{O}_4$ requires 24.0 mL of 0.0100 M KMnO_4 solution to reach the equivalence point, what is the percent by mass of $\text{H}_2\text{C}_2\text{O}_4$ in the sample?

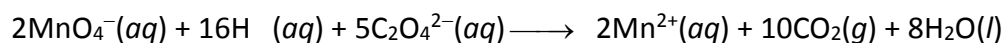
Sol.

(a) The half-reactions are:





We combine the half-reactions to cancel electrons, that is, [2 × equation (i)] [5 × equation (ii)]



(b) We can calculate the moles of KMnO_4 from the molarity and volume of solution.

$$24.0 \text{ mL } \text{KMnO}_4 \times \frac{0.0100 \text{ mol } \text{KMnO}_4}{1000 \text{ mL soln}} = 2.40 \times 10^{-4} \text{ mol } \text{KMnO}_4$$

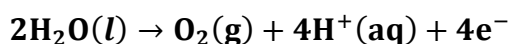
We can calculate the mass of oxalic acid from the stoichiometry of the balanced equation. The mole ratio between oxalate ion and permanganate ion is 5:2.

$$(2.40 \times 10^{-4} \text{ mol } \text{KMnO}_4) \times \frac{5 \text{ mol } \text{H}_2\text{C}_2\text{O}_4}{2 \text{ mol } \text{KMnO}_4} \times \frac{90.04 \text{ g } \text{H}_2\text{C}_2\text{O}_4}{1 \text{ mol } \text{H}_2\text{C}_2\text{O}_4} = 0.0540 \text{ g } \text{H}_2\text{C}_2\text{O}_4$$

Finally, the percent by mass of oxalic acid in the sample is:

$$\% \text{ oxalic acid} = \frac{0.0540 \text{ g}}{1.00 \text{ g}} \times 100\% = \mathbf{5.40\%}$$

3. (5%) One of the half-reactions for the electrolysis of water is



If 0.076 L of O_2 is collected at 25°C and 755 mmHg, how many moles of electrons had to pass through the solution?

Sol.

Find the amount of oxygen using the ideal gas equation

$$n = \frac{PV}{RT} = \frac{\left(755 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right)(0.076 \text{ L})}{(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1})(298 \text{ K})} = 3.1 \times 10^{-3} \text{ mol } \text{O}_2$$

Since the half-reaction shows that one mole of oxygen requires four faradays of

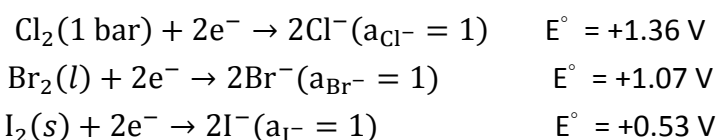
electric charge,
we write

$$(3.1 \times 10^{-3} \text{ mol O}_2) \times \frac{4 F}{1 \text{ mol O}_2} = \mathbf{0.012 F}$$

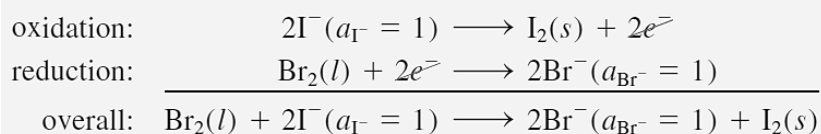
4. (5%) Predict what will happen if molecular bromine (Br₂) is added to a solution containing NaCl and NaI at 25°C. Assume all the species are in their standard states.

Sol.

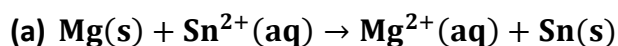
From Table 13.1, we write the standard reduction potentials as follows:



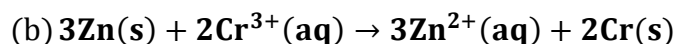
Applying the diagonal rule, we see that Br₂ will oxidize I₂ but will not oxidize Cl₂. Therefore, the only redox reaction that will occur appreciably under standard-state conditions is



5. (10%) Calculate E°, E, and ΔG for the following cell reactions. Assume ideal behavior.



$$[\text{Mg}^{2+}] = 0.045\text{M}, [\text{Sn}^{2+}] = 0.035\text{M}$$



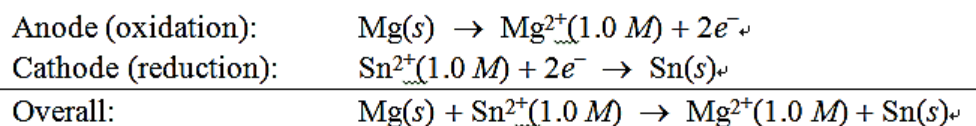
$$[\text{Cr}^{3+}] = 0.010\text{M}, [\text{Zn}^{2+}] = 0.0085\text{M}$$

Sol.

Strategy: The standard emf (E°) can be calculated using the standard reduction potentials in Table 13.1 of the text. Because the reactions are not run under standard-state conditions (concentrations are not 1 M), we need Nernst's equation [Equation 13.10 of the text] to calculate the emf (E) of a hypothetical galvanic cell. Remember that solids do not appear in the reaction quotient (Q) term in the Nernst equation. We can calculate ΔG

from E using Equation 13.2 of the text: $\Delta G = -nFE_{\text{cell}}$.

(a) The half-cell reactions are:



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = E_{\text{Sn}^{2+}/\text{Sn}}^{\circ} - E_{\text{Mg}^{2+}/\text{Mg}}^{\circ}$$

$$E_{\text{cell}}^{\circ} = -0.14\text{ V} - (-2.37\text{ V}) = \mathbf{2.23\text{ V}}$$

From Equation (13.10) of the text, we write:

$$E = E^{\circ} - \frac{0.0257\text{ V}}{n} \ln Q$$

$$E = E^{\circ} - \frac{0.0257\text{ V}}{n} \ln \frac{[\text{Mg}^{2+}]}{[\text{Sn}^{2+}]}$$

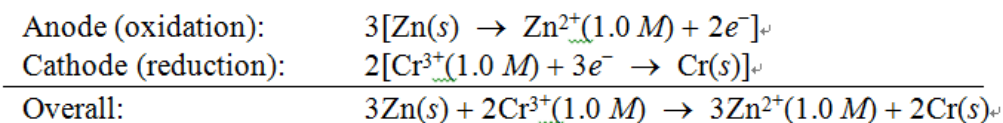
$$E = 2.23\text{ V} - \frac{0.0257\text{ V}}{2} \ln \frac{0.045}{0.035} = \mathbf{2.23\text{ V}}$$

We can now find the free energy change at the given concentrations using Equation 13.2 of the text. Note that in this reaction, $n = 2$.

$$\Delta G = -nFE_{\text{cell}}$$

$$\Delta G = -(2)(96500\text{ J V}^{-1}\text{ mol}^{-1})(2.23\text{ V}) = \mathbf{-430\text{ kJ mol}^{-1}}$$

(b) The half-cell reactions are:



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = E_{\text{Cr}^{3+}/\text{Cr}}^{\circ} - E_{\text{Zn}^{2+}/\text{Zn}}^{\circ}$$

$$E_{\text{cell}}^{\circ} = -0.74\text{ V} - (-0.76\text{ V}) = \mathbf{0.02\text{ V}}$$

From Equation (13.10) of the text, we write:

$$E = E^{\circ} - \frac{0.0257\text{ V}}{n} \ln Q$$

$$E = E^{\circ} - \frac{0.0257\text{ V}}{n} \ln \frac{[\text{Zn}^{2+}]^3}{[\text{Cr}^{3+}]^2}$$

$$E = 0.02 \text{ V} - \frac{0.0257 \text{ V}}{6} \ln \frac{(0.0085)^3}{(0.010)^2} = \mathbf{0.04 \text{ V}}$$

We can now find the free energy change at the given concentrations using Equation 13.2 of the text. Note that in this reaction, $n = 6$.

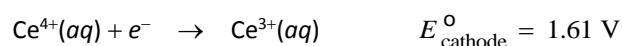
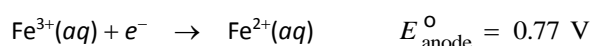
$$\Delta G = -nFE_{\text{cell}}$$

$$\Delta G = -(6)(96500 \text{ J V}^{-1} \text{ mol}^{-1})(0.04 \text{ V}) = \mathbf{-23 \text{ kJ mol}^{-1}}$$

- 6. (10%) Under standard-state conditions, what spontaneous reaction will occur in aqueous solution among the ions Ce^{4+} , Ce^{3+} , Fe^{3+} , and Fe^{2+} ? Calculate ΔG° and K_c for the reaction.**

Sol.

The half-reactions are:



Thus, Ce^{4+} will oxidize Fe^{2+} to Fe^{3+} ; this makes the $\text{Fe}^{2+}/\text{Fe}^{3+}$ half-reaction the anode. The standard cell emf is found using Equation (13.1) of the text.

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ = 1.61 \text{ V} - 0.77 \text{ V} = 0.84 \text{ V}$$

The values of ΔG° and K_c are found using Equations 13.3 and 13.5 of the text.

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ = -(1)(96500 \text{ J V}^{-1} \text{ mol}^{-1})(0.84 \text{ V}) = \mathbf{-81 \text{ kJ mol}^{-1}}$$

$$\ln K = \frac{nE_{\text{cell}}^\circ}{0.0257 \text{ V}}$$

$$K_c = e^{\frac{nE_{\text{cell}}^\circ}{0.0257 \text{ V}}} = e^{\frac{(1)(0.84 \text{ V})}{0.0257 \text{ V}}} = \mathbf{2 \times 10^{14}}$$

- 7. (10%) A galvanic cell consists of a silver electrode in contact with 346 mL of 0.100 M AgNO_3 solution and a magnesium electrode in contact with 288 mL of 0.100 M $\text{Mg}(\text{NO}_3)_2$ solution. (a) Calculate E for the cell at 25°C . (b) A current is drawn from the cell until 1.20 g of silver has been deposited at the silver electrode. Calculate E for the cell at this stage of operation. (molar mass of**

silver = 107.9 g mol⁻¹)

Sol.

- (a) If this were a standard cell, the concentrations would all be 1.00 M, and the voltage would just be the standard emf calculated from Table 13.1 of the text. Since cell emf's depend on the concentrations of the reactants and products, we must use the Nernst equation [Equation (13.10) of the text] to find the emf of a nonstandard cell.

$$E = E^\circ - \frac{0.0257 \text{ V}}{n} \ln Q$$
$$E = 3.17 \text{ V} - \frac{0.0257 \text{ V}}{2} \ln \frac{[\text{Mg}^{2+}]}{[\text{Ag}^+]^2}$$
$$E = 3.17 \text{ V} - \frac{0.0257 \text{ V}}{2} \ln \frac{0.10}{[0.10]^2}$$
$$E = \mathbf{3.14 \text{ V}}$$

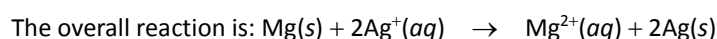
- (b) First we calculate the concentration of silver ion remaining in solution after the deposition of 1.20 g of silver metal

Ag originally in solution: $\frac{0.100 \text{ mol Ag}^+}{1 \text{ L}} \times 0.346 \text{ L} = 3.46 \times 10^{-2} \text{ mol Ag}^+$

Ag deposited: $1.20 \text{ g Ag} \times \frac{1 \text{ mol}}{107.9 \text{ g}} = 1.11 \times 10^{-2} \text{ mol Ag}$

Ag remaining in solution: $(3.46 \times 10^{-2} \text{ mol Ag}) - (1.11 \times 10^{-2} \text{ mol Ag}) = 2.35 \times 10^{-2} \text{ mol Ag}$

$$[\text{Ag}^+] = \frac{2.35 \times 10^{-2} \text{ mol}}{0.346 \text{ L}} = 6.79 \times 10^{-2} \text{ M}$$



We use the balanced equation to find the amount of magnesium metal suffering oxidation and dissolving.

$$(1.11 \times 10^{-2} \text{ mol Ag}) \times \frac{1 \text{ mol Mg}}{2 \text{ mol Ag}} = 5.55 \times 10^{-3} \text{ mol Mg}$$

The amount of magnesium originally in solution was

$$0.288 \text{ L} \times \frac{0.100 \text{ mol}}{1 \text{ L}} = 2.88 \times 10^{-2} \text{ mol}$$

The new magnesium ion concentration is:

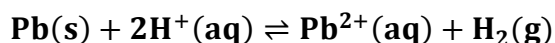
$$\frac{[(5.55 \times 10^{-3}) + (2.88 \times 10^{-2})] \text{ mol}}{0.288 \text{ L}} = 0.119 \text{ M}$$

The new cell emf is:

$$E = E^\circ - \frac{0.0257 \text{ V}}{n} \ln Q$$

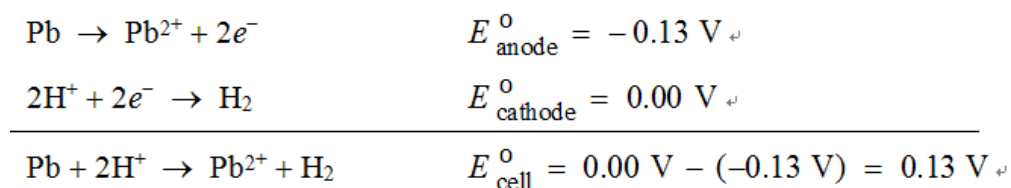
$$E = 3.17 \text{ V} - \frac{0.0257 \text{ V}}{2} \ln \frac{0.119}{(6.79 \times 10^{-2})^2} = 3.13 \text{ V}$$

8. (10%) Calculate the pressure of H_2 (in bar) required to maintain equilibrium with respect to the following reaction at 25°C :



given that $[\text{Pb}^{2+}] = 0.035 \text{ M}$ and the solution is buffered at pH 1.60.

sol.



$$\text{pH} = 1.60$$

$$[\text{H}^+] = 10^{-1.60} = 0.025 \text{ M}$$

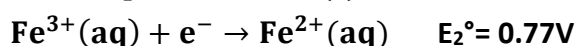
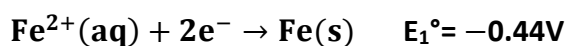
$$E = E^\circ - \frac{RT}{nF} \ln \frac{[\text{Pb}^{2+}]P_{\text{H}_2}}{[\text{H}^+]^2}$$

$$0 = 0.13 - \frac{0.0257 \text{ V}}{2} \ln \frac{(0.035)P_{\text{H}_2}}{0.025^2}$$

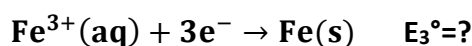
$$\frac{0.26}{0.0257} = \ln \frac{(0.035)P_{\text{H}_2}}{0.025^2}$$

$$P_{\text{H}_2} = 4.4 \times 10^2 \text{ atm}$$

9. (5%) Based on the following standard reduction potentials:



calculate the standard reduction potential for the half-reaction



Sol.

It might appear that because the sum of the first two half-reactions yields the third half-reaction, E_3° is given by $E_1^\circ + E_2^\circ = 0.33 \text{ V}$. This is **not** the case, however, because emf is not an extensive property. Because the number of electrons in the first two half-reactions are not the same we cannot set $E_3^\circ = E_1^\circ + E_2^\circ$. On the other hand, the Gibbs energy is an extensive property, so we can add the separate Gibbs energy changes to obtain the overall Gibbs energy change.

$$\Delta G_3^\circ = \Delta G_1^\circ + \Delta G_2^\circ$$

Substituting the relationship $\Delta G^\circ = -nFE^\circ$, we obtain

$$n_3 FE_3^\circ = n_1 FE_1^\circ + n_2 FE_2^\circ$$

$$E_3^\circ = \frac{n_1 E_1^\circ + n_2 E_2^\circ}{n_3}$$

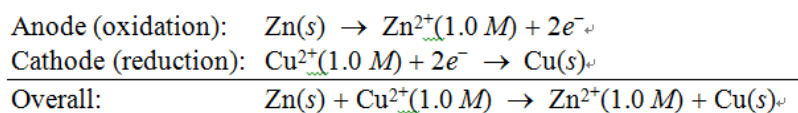
$n_1 = 2$, $n_2 = 1$, and $n_3 = 3$.

$$E_3^\circ = \frac{(2)(-0.44 \text{ V}) + (1)(0.77 \text{ V})}{3} = -0.037 \text{ V}$$

10. (10%) Calculate the emf of the Daniell cell at 25°C when the concentrations of CuSO₄ and ZnSO₄ are 0.50 M and 0.10 M respectively. What would be the emf if activities were used instead of concentrations? (The γ_{\pm} for CuSO₄ and ZnSO₄ at their respective concentrations are 0.068 and 0.15, respectively.)

Sol.

A schematic of the Daniell cell is shown in Figure 13.3. We need to start by determining the standard potential for this cell.



$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ = E_{\text{Cu}^{2+}/\text{Cu}}^\circ - E_{\text{Zn}^{2+}/\text{Zn}}^\circ$$

$$E_{\text{cell}}^\circ = 0.34 \text{ V} - (-0.76 \text{ V}) = 1.10 \text{ V}$$

Now we can use the Nernst equation [see Equation 13.9 in the text] to calculate the potential at non-standard concentrations.

$$E = E^\circ - \frac{0.0257 \text{ V}}{n} \ln Q$$

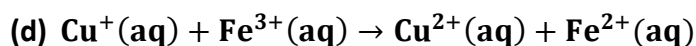
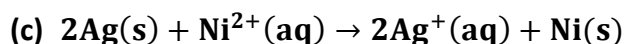
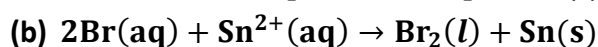
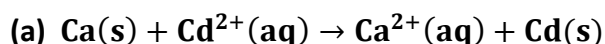
$$E = E^\circ - \frac{0.0257 \text{ V}}{n} \ln \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$E = 1.10 \text{ V} - \frac{0.0257 \text{ V}}{2} \ln \left(\frac{0.10 \text{ M}}{0.50 \text{ M}} \right) = 1.12 \text{ V}$$

If we use activities instead of concentrations, we get the following result:

$$E = 1.10 \text{ V} - \frac{0.0257 \text{ V}}{2} \ln \left(\frac{0.15 \text{ M}}{0.068 \text{ M}} \right) = 1.09 \text{ V}$$

11. (10%) Predict whether the following reactions would occur spontaneously in aqueous solution at 25°C. Assume all species are in their standard states.



Sol.

Strategy: E°_{cell} is *positive* for a spontaneous reaction. In each case, we can calculate the standard cell emf from the potentials for the two half-reactions.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

Solution:

(a) $E^\circ = -0.40 \text{ V} - (-2.87 \text{ V}) = 2.47 \text{ V}$. The reaction is spontaneous.

(b) $E^\circ = -0.14 \text{ V} - 1.07 \text{ V} = -1.21 \text{ V}$. The reaction is not spontaneous.

(c) $E^\circ = -0.25 \text{ V} - 0.80 \text{ V} = -1.05 \text{ V}$. The reaction is not spontaneous.

(d) $E^\circ = 0.77 \text{ V} - 0.15 \text{ V} = 0.62 \text{ V}$. The reaction is spontaneous.

12. (10%) Consider a Daniell cell operating under non-standard-state conditions.

Suppose that the cell reaction is multiplied by 2. What effect does this have on each of the following quantities in the Nernst equation: (a) E° , (b) E , (c) Q , (d) $\ln Q$, and (e) n ?

Sol.

(a) unchanged

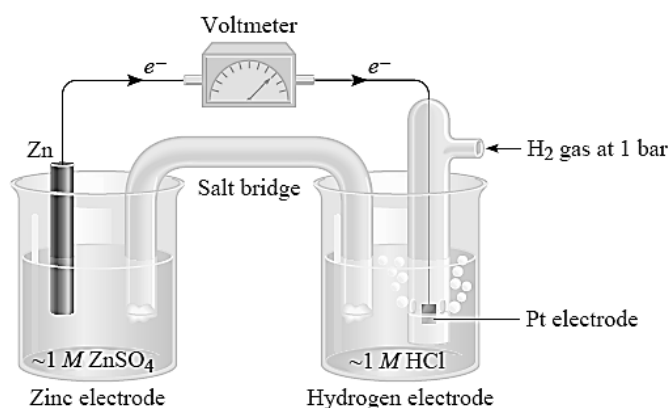
(b) unchanged

(c) squared

(d) doubled

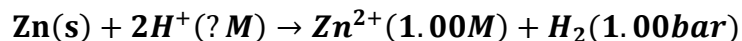
(e) doubled

13. (5%) Consider the galvanic cell as following figure. In a certain experiment, the emf (E) of the cell is found to be 0.54 V at 25°C. Suppose that $[\text{Zn}^{2+}] = 1.00 \times 10^{-4} \text{ M}$ and $P_{\text{H}_2} = 1.00 \text{ bar}$. Calculate the molar concentration of H^+ , assuming ideal solution behavior.



Sol.

Strategy The equation that relates standard emf and nonstandard emf is the Nernst equation. Assuming ideal solution behavior, the standard states of Zn^{2+} and H^+ are 1 M solutions. The overall cell reaction is



Given the emf of the cell (E), we apply the Nernst equation to solve for $[\text{H}^+]$. Note that 2 moles of electrons are transferred per mole of reaction, that is, $n = 2$.

Solution As we saw earlier (page 675), the standard emf (E°) for the cell is 0.76 V. Because the system is at 25°C (298 K) we can use the form of the Nernst equation given in Equation 13.5:

$$\begin{aligned}
\mathcal{E} &= \mathcal{E}^\circ - \frac{0.0257 \text{ V}}{n} \ln Q \\
&= \mathcal{E}^\circ - \frac{0.0257 \text{ V}}{n} \ln \frac{a_{\text{Zn}^{2+}} a_{\text{H}_2}}{a_{\text{H}^+}^2} \\
&\approx \mathcal{E}^\circ - \frac{0.0257 \text{ V}}{n} \ln \frac{[\text{Zn}^{2+}] P_{\text{H}_2}}{[\text{H}^+]^2} \\
0.54 \text{ V} &= 0.76 \text{ V} - \frac{0.0257 \text{ V}}{2} \ln \frac{(1.00 \times 10^{-4})(1.00)}{[\text{H}^+]^2} \\
-0.34 \text{ V} &= \frac{0.0257 \text{ V}}{2} \ln \frac{1}{[\text{H}^+]^2} \\
26.4 &= \ln \frac{1}{[\text{H}^+]^2} \\
e^{26.4} &= \frac{1}{[\text{H}^+]^2} \\
[\text{H}^+] &= \sqrt{\frac{1}{3.1 \times 10^{11}}} = 1.8 \times 10^{-6} M
\end{aligned}$$

Table 13.1 Standard Reduction Potentials at 25°C*

	Half-Reaction	$E^\circ(\text{V})$
↑ Increasing strength as oxidizing agent	$\text{F}_2(\text{g}) + 2\text{e}^- \longrightarrow 2\text{F}^-(\text{aq})$	+2.87
	$\text{O}_3(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{O}_2(\text{g}) + \text{H}_2\text{O}$	+2.07
	$\text{Co}^{3+}(\text{aq}) + \text{e}^- \longrightarrow \text{Co}^{2+}(\text{aq})$	+1.82
	$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow 2\text{H}_2\text{O}$	+1.77
	$\text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{e}^- \longrightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}$	+1.70
	$\text{Ce}^{4+}(\text{aq}) + \text{e}^- \longrightarrow \text{Ce}^{3+}(\text{aq})$	+1.61
	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \longrightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}$	+1.51
	$\text{Au}^{3+}(\text{aq}) + 3\text{e}^- \longrightarrow \text{Au}(\text{s})$	+1.50
	$\text{Cl}_2(\text{g}) + 2\text{e}^- \longrightarrow 2\text{Cl}^-(\text{aq})$	+1.36
	$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \longrightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}$	+1.33
	$\text{MnO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Mn}^{2+}(\text{aq}) + 2\text{H}_2\text{O}$	+1.23
	$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \longrightarrow 2\text{H}_2\text{O}$	+1.23
	$\text{Br}_2(\text{l}) + 2\text{e}^- \longrightarrow 2\text{Br}^-(\text{aq})$	+1.07
	$\text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3\text{e}^- \longrightarrow \text{NO}(\text{g}) + 2\text{H}_2\text{O}$	+0.96
	$2\text{Hg}_2^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Hg}_2^{2+}(\text{aq})$	+0.92
	$\text{Hg}_2^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow 2\text{Hg}(\text{l})$	+0.85
	$\text{Ag}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Ag}(\text{s})$	+0.80
	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \longrightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
	$\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{H}_2\text{O}_2(\text{aq})$	+0.68
	$\text{MnO}_4^-(\text{aq}) + 2\text{H}_2\text{O} + 3\text{e}^- \longrightarrow \text{MnO}_2(\text{s}) + 4\text{OH}^-(\text{aq})$	+0.59
	$\text{I}_2(\text{s}) + 2\text{e}^- \longrightarrow 2\text{I}^-(\text{aq})$	+0.53
	$\text{O}_2(\text{g}) + 2\text{H}_2\text{O} + 4\text{e}^- \longrightarrow 4\text{OH}^-(\text{aq})$	+0.40
	$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cu}(\text{s})$	+0.34
	$\text{AgCl}(\text{s}) + \text{e}^- \longrightarrow \text{Ag}(\text{s}) + \text{Cl}^-(\text{aq})$	+0.22
	$\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}$	+0.20
	$\text{Cu}^{2+}(\text{aq}) + \text{e}^- \longrightarrow \text{Cu}^+(\text{aq})$	+0.15
	$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Sn}^{2+}(\text{aq})$	+0.13
	$2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{H}_2(\text{g})$	0.00
	$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Pb}(\text{s})$	-0.13
	$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Sn}(\text{s})$	-0.14
	$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Ni}(\text{s})$	-0.25
	$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Co}(\text{s})$	-0.28
	$\text{PbSO}_4(\text{s}) + 2\text{e}^- \longrightarrow \text{Pb}(\text{s}) + \text{SO}_4^{2-}(\text{aq})$	-0.31
	$\text{Cd}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cd}(\text{s})$	-0.40
	$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Fe}(\text{s})$	-0.44
	$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^- \longrightarrow \text{Cr}(\text{s})$	-0.74
	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Zn}(\text{s})$	-0.76
	$2\text{H}_2\text{O} + 2\text{e}^- \longrightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$	-0.83
	$\text{Mn}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Mn}(\text{s})$	-1.18
	$\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \longrightarrow \text{Al}(\text{s})$	-1.66
	$\text{Be}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Be}(\text{s})$	-1.85
	$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Mg}(\text{s})$	-2.37
	$\text{Na}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Na}(\text{s})$	-2.71
	$\text{Ca}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Ca}(\text{s})$	-2.87
	$\text{Sr}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Sr}(\text{s})$	-2.89
	$\text{Ba}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Ba}(\text{s})$	-2.90
	$\text{K}^+(\text{aq}) + \text{e}^- \longrightarrow \text{K}(\text{s})$	-2.93
	$\text{Li}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Li}(\text{s})$	-3.05
		↓ Increasing strength as reducing agent

*For all half-reactions the activity is unity for dissolved species (~1 M) and the pressure is 1 bar for gases. These are the standard-state values.