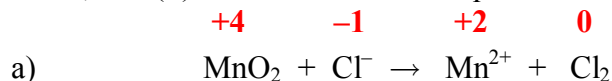


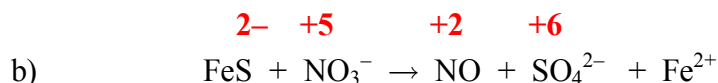
**REVIEW QUESTIONS**

## Chapter 19

1. For each of the following unbalanced equations, (i) write the half-reactions for oxidation and reduction, and (ii) balance the overall equation in acidic solution using the half-reaction method.

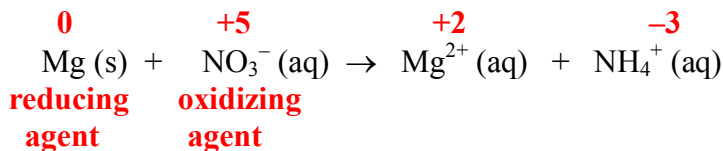


Step	Action	Equations
1	Half-Reactions	$2 \text{Cl}^- \rightarrow \text{Cl}_2 + 2 \text{e}^-$ (oxid) $\text{MnO}_2 + 2 \text{e}^- \rightarrow \text{Mn}^{2+}$ (red)
2	Balance atoms: (except H and O)	No action necessary
3	Balance O: (with H <sub>2</sub> O)	$2 \text{Cl}^- \rightarrow \text{Cl}_2 + 2 \text{e}^-$ $\text{MnO}_2 + 2 \text{e}^- \rightarrow \text{Mn}^{2+} + 2 \text{H}_2\text{O}$
4	Balance H: (with H <sup>+</sup> )	$2 \text{Cl}^- \rightarrow \text{Cl}_2 + 2 \text{e}^-$ $4 \text{H}^+ + \text{MnO}_2 + 2 \text{e}^- \rightarrow \text{Mn}^{2+} + 2 \text{H}_2\text{O}$
5	Balance electrons: (multiply by factor)	No action necessary
6	Combine and simplify	
		$4 \text{H}^+ + \text{MnO}_2 + 2 \text{Cl}^- \rightarrow \text{Cl}_2 + \text{Mn}^{2+} + 2 \text{H}_2\text{O}$



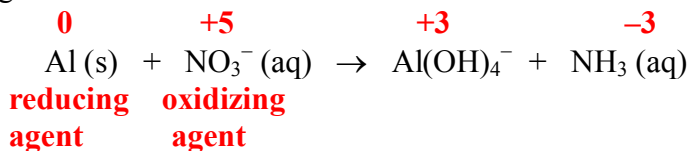
Step	Action	Equations
1	Half-Reactions	$\text{FeS} \rightarrow \text{SO}_4^{2-} + \text{Fe}^{2+} + 8 \text{e}^-$ (oxid) $\text{NO}_3^- + 3 \text{e}^- \rightarrow \text{NO}$ (red)
2	Balance atoms: (except H and O)	No action necessary
3	Balance O: (with H <sub>2</sub> O)	$4 \text{H}_2\text{O} + \text{FeS} \rightarrow \text{SO}_4^{2-} + \text{Fe}^{2+} + 8 \text{e}^-$ $\text{NO}_3^- + 3 \text{e}^- \rightarrow \text{NO} + 2 \text{H}_2\text{O}$
4	Balance H: (with H <sup>+</sup> )	$4 \text{H}_2\text{O} + \text{FeS} \rightarrow \text{SO}_4^{2-} + \text{Fe}^{2+} + 8 \text{e}^- + 8 \text{H}^+$ $4 \text{H}^+ + \text{NO}_3^- + 3 \text{e}^- \rightarrow \text{NO} + 2 \text{H}_2\text{O}$
5	Balance electrons: (multiply by factor)	$12 \text{H}_2\text{O} + 3 \text{FeS} \rightarrow 3 \text{SO}_4^{2-} + 3 \text{Fe}^{2+} + 24 \text{e}^- + 24 \text{H}^+$ $32 \text{H}^+ + 8 \text{NO}_3^- + 24 \text{e}^- \rightarrow 8 \text{NO} + 16 \text{H}_2\text{O}$
6	Combine and simplify	
		$3 \text{FeS} + 8 \text{NO}_3^- + 8 \text{H}^+ \rightarrow 3 \text{SO}_4^{2-} + 3 \text{Fe}^{2+} + 8 \text{NO} + 4 \text{H}_2\text{O}$

2. Balance the following redox reaction in acidic solution, and determine the oxidizing and reducing agents.



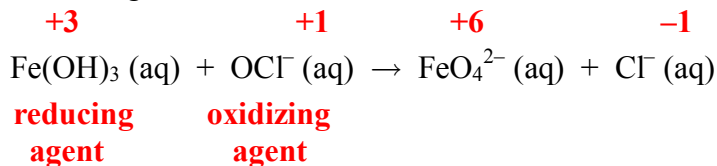
Step	Action	Equations
1	Half-Reactions	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^-$ (oxid) $\text{NO}_3^- + 8 \text{e}^- \rightarrow \text{NH}_4^+$ (red)
2	Balance atoms: (except H and O)	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^-$ $\text{NO}_3^- + 8 \text{e}^- \rightarrow \text{NH}_4^+$
3	Balance O: (with H <sub>2</sub> O)	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^-$ $\text{NO}_3^- + 8 \text{e}^- \rightarrow \text{NH}_4^+ + 3 \text{H}_2\text{O}$
4	Balance H: (with H <sup>+</sup> )	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^-$ $\text{NO}_3^- + 10 \text{H}^+ + 8 \text{e}^- \rightarrow \text{NH}_4^+ + 3 \text{H}_2\text{O}$
5	Balance electrons: (multiply by factor)	$4 \text{Mg} \rightarrow 4 \text{Mg}^{2+} + 8 \text{e}^-$ $\text{NO}_3^- + 10 \text{H}^+ + 8 \text{e}^- \rightarrow \text{NH}_4^+ + 3 \text{H}_2\text{O}$
6		Combine and simplify $4 \text{Mg} + \text{NO}_3^- + 10 \text{H}^+ \rightarrow 4 \text{Mg}^{2+} + \text{NH}_4^+ + 3 \text{H}_2\text{O}$

3. Balance the following redox reaction in basic solution, and determine the oxidizing and reducing agents.



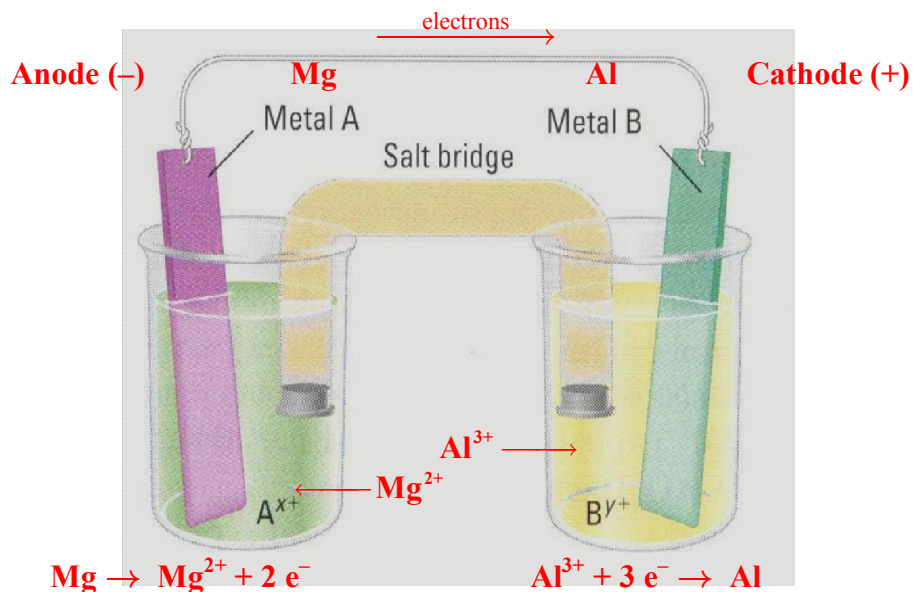
Step	Action	Equations
1	Half-Reactions	$\text{Al} \rightarrow \text{Al(OH)}_4^- + 3 \text{e}^-$ (oxid) $\text{NO}_3^- + 8 \text{e}^- \rightarrow \text{NH}_3$ (red)
2	Balance atoms: (except H and O)	$\text{Al} \rightarrow \text{Al(OH)}_4^- + 3 \text{e}^-$ $\text{NO}_3^- + 8 \text{e}^- \rightarrow \text{NH}_3$
3	Balance O: (with OH <sup>-</sup> and H <sub>2</sub> O)	$\text{Al} + 4 \text{OH}^- \rightarrow \text{Al(OH)}_4^- + 3 \text{e}^-$ $\text{NO}_3^- + 3 \text{H}_2\text{O} + 8 \text{e}^- \rightarrow \text{NH}_3 + 6 \text{OH}^-$
4	Balance H: (with H <sub>2</sub> O and OH <sup>-</sup> )	$\text{Al} + 4 \text{OH}^- \rightarrow \text{Al(OH)}_4^- + 3 \text{e}^-$ $\text{NO}_3^- + 6 \text{H}_2\text{O} + 8 \text{e}^- \rightarrow \text{NH}_3 + 9 \text{OH}^-$
5	Balance electrons: (multiply by factor)	$8 \text{Al} + 32 \text{OH}^- \rightarrow 8 \text{Al(OH)}_4^- + 24 \text{e}^-$ $3 \text{NO}_3^- + 18 \text{H}_2\text{O} + 24 \text{e}^- \rightarrow 3 \text{NH}_3 + 27 \text{OH}^-$
6		Combine and simplify $8 \text{Al} + 5 \text{OH}^- + 3 \text{NO}_3^- + 18 \text{H}_2\text{O} \rightarrow 8 \text{Al(OH)}_4^- + 3 \text{NH}_3$

4. Balance the following redox reaction in basic solution, and determine the oxidizing and reducing agents.

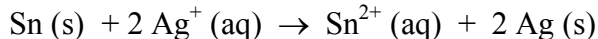


Step	Action	Equations
1	Half-Reactions	$\text{Fe}(\text{OH})_3 \rightarrow \text{FeO}_4^{2-} + 3 \text{e}^- \text{ (oxid)}$ $\text{OCl}^- + 2 \text{e}^- \rightarrow \text{Cl}^- \text{ (red)}$
2	Balance atoms: (except H and O)	No action necessary
3	Balance O: (with $\text{OH}^-$ and $\text{H}_2\text{O}$ )	$2 \text{OH}^- + \text{Fe}(\text{OH})_3 \rightarrow \text{FeO}_4^{2-} + 3 \text{e}^- + \text{H}_2\text{O}$ $\text{H}_2\text{O} + \text{OCl}^- + 2 \text{e}^- \rightarrow \text{Cl}^- + 2 \text{OH}^-$
4	Balance H: (with $\text{H}_2\text{O}$ and $\text{OH}^-$ )	$5 \text{OH}^- + \text{Fe}(\text{OH})_3 \rightarrow \text{FeO}_4^{2-} + 3 \text{e}^- + 4 \text{H}_2\text{O}$ $\text{H}_2\text{O} + \text{OCl}^- + 2 \text{e}^- \rightarrow \text{Cl}^- + 2 \text{OH}^-$
5	Balance electrons: (multiply by factor)	$10 \text{OH}^- + 2 \text{Fe}(\text{OH})_3 \rightarrow 2 \text{FeO}_4^{2-} + 6 \text{e}^- + 8 \text{H}_2\text{O}$ $3 \text{H}_2\text{O} + 3 \text{OCl}^- + 6 \text{e}^- \rightarrow 3 \text{Cl}^- + 6 \text{OH}^-$
6		<p>Combine and simplify</p> $2 \text{Fe}(\text{OH})_3 + 3 \text{OCl}^- + 4 \text{OH}^- \rightarrow 2 \text{FeO}_4^{2-} + 3 \text{Cl}^- + 5 \text{H}_2\text{O}$

5. The diagram below shows a voltaic cell with the anode on the left side and the cathode on the right side. Given that this is a magnesium and aluminum cell,
- identify metals A and B,
  - identify solutions A and B,
  - write half-reactions for each electrode,
  - direction of electron flow,
  - the polarities of the anode and the cathode,
  - calculate the cell potential, and
  - write a shorthand cell notation.



6. A voltaic cell employs the reaction:



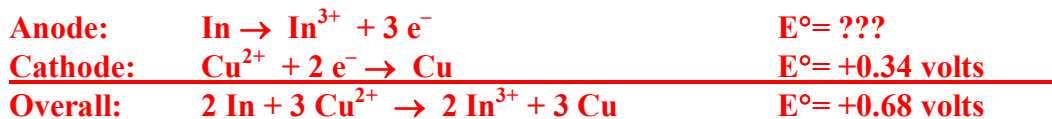
Calculate the voltage produced by this reaction under standard conditions at 25°C. (Use Table 20.1 in your textbook for standard reduction potentials).



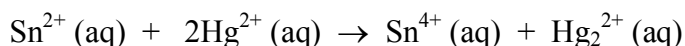
7. The standard voltage ( $\mathcal{E}^\circ$ ) for the voltaic cell shown below is +0.68 volts.



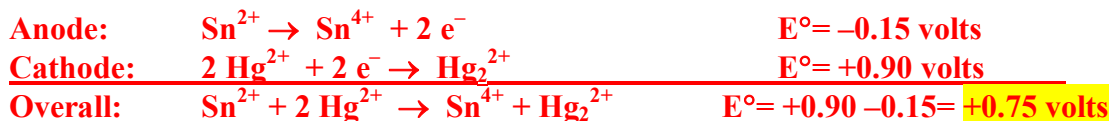
Determine the standard reduction potential for:  $\text{In}^{3+} + 3 \text{e}^- \rightarrow \text{In}$



8. A voltaic cell used the reaction shown below:



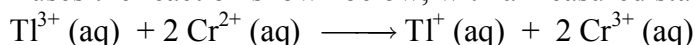
a) Calculate the voltage for this reaction under standard conditions.  
(Use Table 20.1 in your textbook for standard reduction potentials)



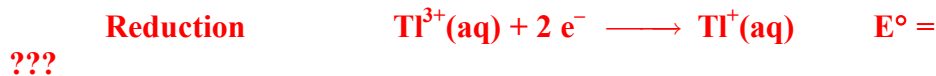
b) Calculate  $\Delta G^\circ$  for this reaction.

$$\Delta G^\circ = w_{\text{max}} = -nFE^\circ = -2 \text{ mole}^- (96500 \text{ C/mole}^-)(0.75 \text{ volts}) = -1.4 \times 10^5 \text{ J}$$

9. A voltaic cell uses the reaction shown below, with a measured standard cell potential of 1.19 V:



- a) Write the two half-cell reactions.

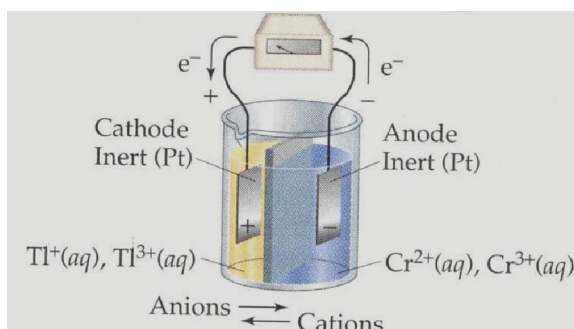


- b) What is the  $E^{\circ}$  for the reduction of  $\text{Tl}^{3+}$ ?

$$E^{\circ}_{\text{cell}} = 0.41 + E^{\circ}(\text{Tl}^{3+}) = 1.19 \text{ V}$$

$$E^{\circ}(\text{Tl}^{3+}) = 1.19 - 0.41 = \mathbf{0.78 \text{ V}}$$

- c) Sketch the voltaic cell, label the anode and the cathode, and indicate the direction of the electron flow.

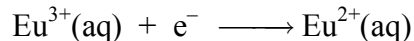


10. For each pair of substances below, use Reduction Potential in your textbook to determine the one that is the stronger oxidizing agent:

**The more positive the reduction potential, the easier the substance is reduced and the stronger the oxidizing agent it will be.**

- a)  $\text{Br}_2(\text{l})$  or  $\text{I}_2(\text{s})$                       (1.065 vs. 0.536)
- b)  $\text{Ag}^{+}(\text{aq})$  or  $\text{AgCl}(\text{s})$                       (0.799 vs. 0.222)
- c)  $\text{Cl}_2(\text{g})$  or  $\text{Ce}^{4+}(\text{aq})$  (1.359 vs. 1.61)
- d)  $\text{Mg}(\text{s})$  or  $\text{K}(\text{s})$                       (-2.37 vs. -2.925)
- e)  $\text{H}_2\text{O}_2(\text{aq})$  or  $\text{Fe}(\text{CN})_6^{4-}(\text{aq})$                       (1.776 vs. 0.36)

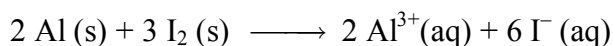
11. The standard cell potential for the reaction shown below is  $-0.43$  V:



Use table of reduction potentials in your textbook to suggest two substances capable of reducing  $\text{Eu}^{3+}$  to  $\text{Eu}^{2+}$ .

**Any species with a standard reduction potential of less than  $-0.43$  can reduce  $\text{Eu}^{3+}$  to  $\text{Eu}^{2+}$ , since it would be a stronger reducing agent than  $\text{Eu}^{3+}$ . Two examples are  $\text{Zn}$  and  $\text{H}_2(\text{g})$ .**

12. The standard cell potential for the reaction shown below at 298 K is 2.20 V:



Calculate the emf generated by this cell when

$$\begin{aligned} [\text{Al}^{3+}] &= 4.0 \times 10^{-3} \text{M} \\ [\text{I}^{-}] &= 0.010 \text{M} \end{aligned}$$

**Writing the two half-reactions:**



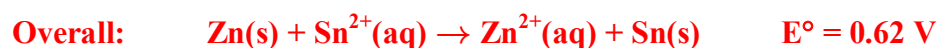
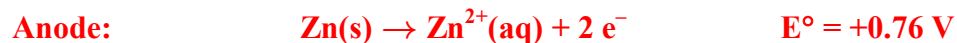
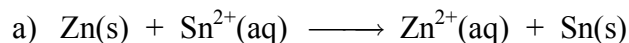
**Using the Nernst equation:**

$$E = E^{\circ} - \frac{0.0592}{n} \log Q \quad n = 6$$

$$E = 2.20 - \frac{0.0592}{6} \log (4.0 \times 10^{-3})^2 (0.010)^6$$

$$E = 2.20 - (-0.166) = 2.37$$

13. Use the standard reduction potentials listed in your textbook to determine the equilibrium constant for each of the following reactions:

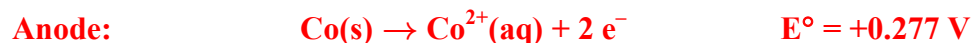
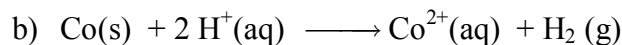


**At equilibrium, Nernst equation can be simplified to:**

$$0 = E^{\circ} - \frac{0.0592}{n} \log K$$

$$\log K = \frac{nE^{\circ}}{0.0592} = \frac{2(0.62)}{0.0592} = 20.95$$

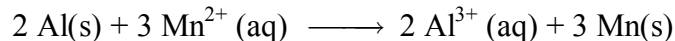
$$K = 10^{20.95} = 8.9 \times 10^{20}$$



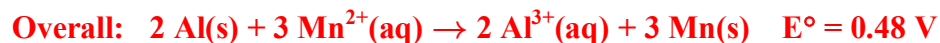
$$\log K = \frac{nE^{\circ}}{0.0592} = \frac{2(0.277)}{0.0592} = 9.36$$

$$K = 10^{9.36} = 2.3 \times 10^9$$

14. A voltaic cell utilizes the reaction shown below at 298 K:



a) Calculate the emf for this cell under standard conditions.



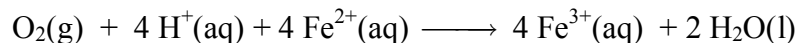
b) Calculate the emf for this cell when  $[\text{Al}^{3+}] = 1.5 \text{ M}$  and  $[\text{Mn}^{2+}] = 0.10 \text{ M}$ .

$$E = E^{\circ} - \frac{0.0592}{n} \log Q \quad n = 6$$

$$E = 0.48 - \frac{0.0592}{6} \log \frac{(1.5)^2}{(0.10)^3}$$

$$E = 0.48 - (0.033) = 0.45 \text{ V}$$

15. Use the standard reduction potentials listed in your textbook to determine the equilibrium constant the following reactions:



$$\log K = \frac{nE^{\circ}}{0.0592} = \frac{4(0.46)}{0.0592} = 31.08$$

$$K = 10^{31.08} = 1.2 \times 10^{31}$$