Lesson Check Answers

1. a redox reaction
2. spontaneous redox reactions within the cell
3. fuel cells, lead storage batteries, and dry cells
4. calcium
5. concentrated sulfuric acid;
   Anode:
   $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$
   Cathode:
   $PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$
6. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$; water is the product. H_2 is oxidized at the anode; O_2 is reduced at the cathode.
7. no reaction

Sample Problems

8. The half-reactions are

Oxidation: $Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^-$

Reduction: $Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$

Writing both half-cells as reductions:

$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s) \qquad E°_{Cr^{3+}} = -0.74\ V$

$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s) \qquad E°_{Zn^{2+}} = -0.76\ V$

$E°_{cell} = E°_{red} - E°_{oxid} = E°_{Zn^{2+}} - E°_{Cr^{3+}}$

$= -0.76\ V - (-0.74\ V) = -0.02\ V$

$E°_{cell} < 0$, so the reaction is not spontaneous.

Another way to calculate: $E^0_{cell} = E^0_{SRP} + E^0_{SOP}$ … negate the oxidation SRP. $E^0_{cell} = E^0_{SRP\ (Zn)} + E^0_{SOP\ (Cr)} =$ -0.76 V + 0.74 V = -0.02 V

9. The half-reactions are

Oxidation: $Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$

Reduction: $Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$

Writing both half-cells as reductions:

$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$ $\quad E°_{Fe^{2+}} = -0.44\ V$

$Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$ $\quad E°_{Co^{2+}} = -0.28\ V$

$E°_{cell} = E°_{red} - E°_{oxid} = E°_{Co^{2+}} - E°_{Fe^{2+}}$

$= -0.28\ V - (-0.44\ V) = +0.16\ V$

$E°_{cell} > 0$, so the reaction is spontaneous.

Another way to calculate: $E^0_{cell} = E^0_{SRP} + E^0_{SOP}$ … negate the oxidation SRP. $E^0_{cell} = E^0_{SRP\ (Co)} + E^0_{SOP\ (Fe)} =$ -0.28 V + 0.44 V = +0.16 V

10. Cu^{2+} is reduced, and Al is oxidized.

Oxidation: $Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$

Reduction: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Writing both half-cells as reductions:

$2[Al(s) \longrightarrow Al^{3+}(aq) + 3e^-]$

$3[Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)]$

$2Al(s) \longrightarrow 2Al^{3+}(aq) + \cancel{6e^-}$

$3Cu^{2+}(aq) + \cancel{6e^-} \longrightarrow 3Cu(s)$

$2Al(s) + 3Cu^{2+}(aq) \longrightarrow 2Al^{3+}(aq) + 3Cu(s)$

11. Ag^+ is reduced, and Cu is oxidized.

The half-reactions are

Oxidation: $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

Reduction: $Ag^+(aq) + e^- \longrightarrow Ag(s)$

Writing both half-cells as reductions:

$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

$2[Ag^+(aq) + e^- \longrightarrow Ag(s)]$

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

$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

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

Notice that copper is oxidized (because it is the more active metal) and will reduce the silver.

12. $E°_{cell} = E°_{red} - E°_{oxid} = E°_{Cu^{2+}} - E°_{Al^{3+}}$

$= 0.34\ V - (-1.66\ V) = +2.00\ V$

Another way to calculate: $E^0_{cell} = E^0_{SRP} + E^0_{SOP}$ … negate the oxidation SRP. $E^0_{cell} = E^0_{SRP\ (Cu)} + E^0_{SOP\ (Al)} = +0.34\ V + 1.66\ V = +2.00\ V$

Aluminum is a more active metal than copper so aluminum will be oxidized and will reduce the copper.

13. $E°_{cell} = E°_{red} - E°_{oxid} = E°_{Ag^+} - E°_{Cu^{2+}}$

$= 0.80\ V - (+0.34\ V) = +0.46\ V$

Another way to calculate: $E^0_{cell} = E^0_{SRP} + E^0_{SOP}$ … negate the oxidation SRP. $E^0_{cell} = E^0_{SRP\ (Ag)} + E^0_{SOP\ (Cu)} = +0.80\ V - 0.34\ V = +0.46\ V$

Copper is a more active metal than silver so copper will be oxidized and will reduce the silver.

Lesson Check Answers

14. Electrical potential of cell results from competition for electrons between two half-cells.

15. by connecting it to a standard hydrogen electrode and measuring the cell potential

16. If the cell potential for a redox reaction is positive, the reaction is spontaneous; if it is negative, the reaction is nonspontaneous.

17. spontaneous; $E^\circ_{cell} = +0.71$ V

18. $2Li(s) + Mg^{2+}(aq) \rightarrow Mg(s) + 2Li^{+}(aq)$; $E^\circ_{cell} = +0.68$

19. There is an positive electrical potential between the negative and positive terminals of the battery. Cell reaction: $Pb + PbO_2 + 2H_2SO_4^{2-} \rightarrow 2PbSO_4 + 2H_2O$; $E^\circ_{cell} = +2.05$ V

17. $2Al^{+3}(aq) + 3Mg^{0}(s) \rightarrow 2Al(s) + 3Mg^{+2}(aq)$

Oxid ½ rxn: $3[Mg^0 \rightarrow Mg^{+2} + 2e^-]$

Red ½ rxn: $2[Al^{+3} + 3e^- \rightarrow Al^0]$

$E^0_{cell} = E^0_{SRP} + E^0_{SOP}$ … negate the oxidation SRP.
$E^0_{cell} = E^0_{SRP\ (Al)} + E^0_{SOP\ (Mg)} = +2.37\text{ V} + (-1.66\text{ V}) = +0.71\text{ V}$
This is a spontaneous reaction.

Magnesium is a more active metal than Aluminum so magnesium will be oxidized and will reduce the aluminum.

18. $2Li(s) + Mg^{+2}(aq) \rightarrow Mg(s) + 2Li^{+}(aq)$

Oxid ½ rxn: $2[Li^0 \rightarrow Li^+ + e^-]$

Red ½ rxn: $Mg^{+2} + 2e^- \rightarrow Mg^0$

$E^0_{cell} = E^0_{SRP} + E^0_{SOP}$ … negate the oxidation SRP.
$E^0_{cell} = E^0_{SRP\ (Mg)} + E^0_{SOP\ (Li)} = -2.37\text{ V} + 3.05\text{ V} = +0.68\text{ V}$
This is a spontaneous reaction.

Lithium is a more active metal than magnesium so lithium will be oxidized and will reduce the magnesium.

20. A voltaic cell uses an electrochemical reaction to produce electrical energy; an electrolytic cell uses electrical energy to bring about a chemical change.

21. to separate elements from compounds and to plate, purify, and refine metals

22. electrolytic cell anode (+); voltaic cell anode (–)

23. Reduction always occurs at the cathode.

24. Anode: $C(s) + 2O^{2-}(l) \rightarrow CO_2(g) + 4e^-$

Cathode: $3e^- + Al^{3+}(l) \rightarrow Al(l)$

Adding cryolite to aluminum oxide results in a mixture that melts at a much lower temperature (1012°C) than pure aluminum oxide (2045°C).

25. Yes, the electrical energy produced in a voltaic cell can do work. An electrolytic cell does not produce free energy unless an electrical current is supplied.