Chapter 7

Page 433, Practice Problems 7.1.1
1. H = +1
   O = -2
2. Cs = +1
   O = -1
3. C = +4
   O = -2
4. Na = +1
   Cr = +6
   O = -2
5. Ba = +2
   H = -1
6. N = -3
   H = +1
7. S = 0
8. Al = +3
   S = +6
   O = -2

Page 434, Quick Check
1. 6 e-
2. nothing (no e- transferred)

Page 435, Quick Check
1. You – oxidized (LEO)
   Partner - reduced (GER)
2. Which partner acted as an "oxidizing agent"?
   me (I got oxidized)
   A "reducing agent"?   my partner (they got reduced)
3. What happened to your "oxidation state"?
   Increased (I was oxidized)
   What about your partner?  Reduced – got smaller (they were reduced)

Page 436, Practice Problems 7.1.2
0                    0            -4 +1
1. C(s) + 2 H2(g) → CH4(g) O.A. = C(s)
   R.A. = H2(g)
   0                    0 +3 -1   0 +2 -1
2. 3 Sr(s) + 2 FeBr3(aq) → 2 Fe(s) + 3 SrBr2(aq)
   O.A. = Fe3+        R.A. = Sr
   +2 -2        +5 -2   +4 -2   0
3. 5 CO(g) + Cl2O4(s) → 5 CO2(g) + Cl2(s)
   O.A. = Cl2O4      R.A. = CO
   -3 +1            0 0
4. 4 PH3(g) → P4(g) + 6 H2(g)
   O.A. and R.A. = PH3
   0                    0 +1 -2  +2 -2 +1 0
5. Ba(s) + 2 H2O(l) → Ba(OH)2(s) + H2(g)
   O.A. = H2O        R.A. = Ba
   0                    0 +1 -1 +3 0 +1 -2 +2 -1 0

Page 438, 7.1 Review Questions
1. Elements that get oxidized (act as reducing agents) for (a) + ions when they react. This
means reducing agents are generally (b) metals. Reducing agents may also be (c)
_negatively_ charged ions. The most active reducing agents likely belong to the (d) alkali
metal family on the periodic table. The most active oxidizing agents must belong to the (e)
halogen family.

2. a) CaI2  b) OF2  c) C6H12O6  d) Rb2O2
   +2 +2  0 -1
   e) S2O32-  f) BeH2  g) BrO-  h) Cl2
   +2 -1 +1 0
3. a) A species that gets reduced/causes other species to be oxidized/gains e-
   b) A species that gets oxidized/causes other species to be reduced/loses e-
   c) Increased electronegativity makes a stronger O.A. and a weaker R.A.
4. For each of the following reactions, indicate the species being oxidized and reduced and
   show the oxidation states above their symbols.
   +1 +5 -2 +1 -1 0
   a) 2 KBrO3(s) → 2 KBr(s) + 3 O2(g)
      Oxidized: KBrO3  Reduced: KBr
   0 +1 +5 -2 +2 +5 -2 0
   b) Sr(s) + 2 CuNO3(aq) → Sr(NO3)2(aq) + 2 Cu(s)
      Ox'd: Sr  Red'd: Cu+
   0 0 +2 -1
   c) 2 F2(g) + O2(g) → 2 OF2(g)
      Oxidized: O2  Reduced: F2
   -3 +1+5 -2 +1 -2 +1 -2
   d) NH4NO3(s) → N2O(g) + 2 H2O(l)
      Ox'd: NH4NO3  Red'd: NH4NO3 (solid)

5. Determine the oxidizing and reducing agent in each of the following reactions. Then
indicate the number of electrons transferred by one atom of the reducing agent.
   0 0 +2 -2
   a) 2 Sn(s) + O2(g) → 2 SnO(s)
      OA: O2  RA: Sn  No e⁻: 2
   0 0 +5 -1
   b) 2 V(s) + 5 I2(g) → 2 Vl5(s)
      OA: I2  RA: V  No e⁻: 5
   0 +1 -1 +2 -1 0
   c) Sr(s) + 2 HCl(aq) → SrCl2(aq) + H2(g)
      OA: H+  RA: Sr  No e⁻: 2
   -8/3 +1 0 +4 -2 +1 -2
   d) C3H8(g) + 5 O2(g) → 3 CO2(g) + 4 H2O(g)
      OA: O2  RA: C3H8  No e⁻: 6⅔

6. a) Write a balanced redox equation (in net ionic form) to show what has occurred in the
beaker over time:  

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

b) What is the oxidizing agent?  \(\text{Cu}^{2+}(aq)\) The reducing agent?  \(\text{Fe(s)}\)

c) How many electrons were transferred in the equation?  \(2e^-\)

7. Give the oxidation state of the underlined element in each species:

a)  \(\underset{-3}{\text{P}}_3\)

b)  \((\text{NH}_4)_2\text{Zr(SO}_4)_2\text{) \ +4}

c)  \(\underset{+3}{\text{Na}_2\text{C}_2\text{O}_4}\)

d)  \(\underset{-2}{\text{N}_2\text{H}_5\text{Cl}}\)

e)  \(\underset{+6}{\text{MnO}_4^{2-}}\)

8.

a) see above

b) Oxidized?  \(\text{Cu metal}\)

Reduced?  \(\text{Ag}^+\) ion

c) Reducing agent?  \(\text{Cu metal}\)

Oxidizing agent?  \(\text{Ag}^+\) ion

d)  \(2e^-\)

a) Alkali metal family (IA)

b) Halogens (VIIA) or (17)

9. a)  \(\text{ClO}_3^-\),  \(\text{ClO}_4^-\),  \(\text{ClO}_7^-\)

b)  \(\text{Cl}_2\),  \(\text{Cl}^-\)

Page 444, Practice Problems 7.2.1

1.  \(\text{Sm} \rightarrow \text{Sm}^{3+} + 3e^-\)  oxidation

2.  \(8e^- + 10 \text{H}^+ + \text{NO}_3^- \rightarrow \text{NH}_4^+ + 3 \text{H}_2\text{O}\)  reduction

3.  \(2e^- + \text{H}_2\text{O} + \text{IO}_4^- \rightarrow \text{IO}_3^- + 20\text{H}^+\)  reduction

4.  \(5 \text{H}_2\text{O} + \text{S}_2\text{O}_5^{2-} \rightarrow 2 \text{SO}_4^{2-} + 10 \text{H}^+ + 8e^-\)  oxidation

5.  \(10e^- + 6 \text{H}_2\text{O} + 2 \text{BrO}_3^- \rightarrow \text{Br}_2 + 12 \text{OH}^-\)  reduction

Page 447, Practice Problems 7.2.2

1.  \(4 \text{Zn} + 8 \text{H}^+ + \text{H}_2\text{AsO}_4 \rightarrow \text{AsH}_3 + 4 \text{H}_2\text{O} + 4 \text{Zn}^{2+}\)

Page 449, Practice Problems 7.2.3

1.  \(2 \text{H}_2\text{O} + \text{XeO}_4^- + 2 \text{OH}^- \rightarrow \text{XeO}_6^{4-} + \text{Xe} + \text{O}_2 + 2 \text{H}_2\text{O}\)

2.  \(6 \text{H}_2\text{O} + 2 \text{BrO}_3^- + 10 \text{Br}^- \rightarrow 6 \text{Br}_2 + 12 \text{OH}^-\)

3.  \(2 \text{H}_2\text{O} + 2 \text{CH}_3\text{COO}^- \rightarrow 2 \text{CH}_4 + 2 \text{CO}_2 + 2 \text{OH}^-\)

Page 452, 7.2 Review Questions

1. e.g. \(2^+\)  \(3^+\)  \(1\) e to products

a. \(3^+\) \(2^-\)  \(5\) e to reactants

b. \(1^-\) \(3^-\)  \(2\) e to reactants

c. \(2^-\) \(4^+\)  \(6\) e to reactants
2. a) $2e^- + 2\text{NO}_3^- + 2\text{H}_2\text{O} \rightarrow \text{N}_2\text{O}_4 + 4\text{OH}^-$ (reduction)  
   b) $2\text{Cr}^{3+} + 7\text{H}_2\text{O} \rightarrow \text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6e^-$ (oxidation)  
   c) $8e^- + \text{ClO}_4^- + 4\text{H}_2\text{O} \rightarrow \text{Cl}^- + 8\text{OH}^-$ (oxidation)  
   d) $\text{S}_2\text{O}_5^{2-} + 3\text{H}_2\text{O} \rightarrow 2\text{SO}_4^{2-} + 6\text{H}^+ + 4e^-$ (oxidation)  
3. a) $14e^- + 16\text{H}^+ + 2\text{ClO}_4^- \rightarrow \text{Cl}_2 + 8\text{H}_2\text{O}$ (reduction)  
   b) $\text{H}_2\text{O} + 2\text{FeO} \rightarrow \text{Fe}_2\text{O}_3 + 2\text{H}^+ + 2e^-$ (oxidation)  
   c) $2\text{H}_2\text{O} + \text{N}_2\text{O}_4 \rightarrow 2\text{NO}_3^- + 4\text{H}^+ + 2e^-$ (oxidation)  
4. a) $4\text{H}_2\text{O} + 4e^- + \text{CrO}_4^{2-} \rightarrow \text{Cr}(	ext{OH})_2 + 6\text{OH}^-$ (reduction)  
   b) $2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_2\text{O}_5^{2-} + 2e^-$ (oxidation)  
   c) $4e^- + 3\text{H}_2\text{O} + \text{IO}_3^- + 2\text{Cl}^- \rightarrow \text{ICl}_2^- + 6\text{OH}^-$ (reduction)  
5. a) $4\text{H}_2\text{O} + 2\text{MnO}_4^- + 3\text{Sn}^{2+} \rightarrow 3\text{Sn}^{4+} + 2\text{MnO}_2 + 8\text{OH}^-$  
   b) $2\text{H}^+ + 2\text{H}_2\text{SO}_3 + 4\text{V}^{2+} \rightarrow 4\text{V}^{3+} + \text{S}_2\text{O}_5^{2-} + 3\text{H}_2\text{O}$  
   c) $3\text{H}_2\text{O} + \text{IO}_3^- + 5\text{I}^- \rightarrow 3\text{I}_2 + 6\text{OH}^-$  
6. a) $6\text{H}^+ + \text{SeO}_3^{2-} + 4\text{F}^- \rightarrow \text{Se} + 2\text{F}_2 + 3\text{H}_2\text{O}$  
   b) $2\text{H}_2\text{O} + 4\text{ReO}_4^- + 3\text{Sb}_2\text{O}_3 \rightarrow 4\text{ReO}_2 + 3\text{Sb}_2\text{O}_3 + 4\text{OH}^-$  
   c) $16\text{H}^+ + 4\text{NO}_3^- + 3\text{Pd} + 18\text{I}^- \rightarrow 3\text{PdI}_6^{2-} + 4\text{NO} + 8\text{H}_2\text{O}$  
   d) $\text{Pb}(	ext{OH})_2^{2-} + \text{BrO}^- \rightarrow \text{PbO}_2^- + \text{Br}^- + \text{H}_2\text{O} + 2\text{OH}^-$  
7. a) $16\text{HCl} + 2\text{K}_2\text{Cr}_2\text{O}_7 + 3\text{CH}_3\text{CH}_2\text{OH} \rightarrow 3\text{CH}_3\text{COOH} + 4\text{CrCl}_3 + 11\text{H}_2\text{O} + 4\text{KCl}$  
   b) $4\text{NaCl} + 2\text{H}_2\text{SO}_4 + \text{MnO}_2 \rightarrow 2\text{Na}_2\text{SO}_4 + \text{MnCl}_2 + 2\text{H}_2\text{O} + \text{Cl}_2$  

Page 460, Practice Problems 7.3.2  
1. no; I$_2$ isn’t a strong enough oxidizing agent to spontaneously react with F$^-$  
2. yes; 3Cu$^{2+}$ + 2Al $\rightarrow$ 2Al$^{3+}$ + 3Cu  
3. yes; Cl$_2$ + 2Br$^-$ $\rightarrow$ Br$_2$ + 2Cl$^-$ (* Question revised from complimentary copy)  
4. no; Al$^{3+}$ isn’t a strong enough oxidizing agent to spontaneously react with Sn

Page 462, Quick Check  
1. Sr  
2. Fe$^{3+}$  
3. I$^-$  
4. Sn$^{2+}$  
5. B$^-$  
6. +0.80 V  
7. +2.87 V

Page 464, Practice Problems 7.3.3  
1. Sn$^{4+}$ + Ni $\rightarrow$ Sn$^{2+}$ + Ni$^{2+}$ (* Question revised from complimentary copy)  
2. 3Cu$^{2+}$ + 2Al $\rightarrow$ 2Al$^{3+}$ + 3Cu  
3. 2Cu$^+$$\rightarrow$ Cu$^{3+}$ + Cu  
4. Cu + Br$_2$ $\rightarrow$ 2Br$^- +$ Cu$^{2+}$

Page 465, Practice Problems 7.3.4  
1. a) $\text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 6\text{Fe}^{3+} + 7\text{H}_2\text{O}$  
   b) 0.0900 M  
   c) 0.183 M  
   d) 1.97 M  
   e) 11.5%  

Page 467, 7.3 Review Questions  
1. a) no  
   b) yes  
2. Either Fe$^{3+}$, Hg$_2$$^{2+}$, Ag$^+$, or Hg$^{2+}$  
3. a) yes; 3Mg + 2Al$^{3+}$$\rightarrow$$3\text{Mg}^{2+} + 2\text{Al}$  
   b) no  
   c) yes; Hg$^{2+} +$ 2Ag $\rightarrow$ Hg + 2Ag$^+$
4.

<table>
<thead>
<tr>
<th>Metals</th>
<th>Non-metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>bottom right of SRP Table</td>
<td>top left of SRP Table</td>
</tr>
<tr>
<td>tend to give electrons</td>
<td>tend to take electrons</td>
</tr>
<tr>
<td>give e⁻ to chemicals above them on the left</td>
<td>take e⁻ from chemicals below them on the right</td>
</tr>
</tbody>
</table>

5. a) yes; Fe + Sn²⁺ → Fe²⁺ + Sn
   b) yes; F₂ + 2Br⁻ → Br₂ + 2F⁻
   c) no
6. a) 2I⁻ + 2Fe³⁺ → I₂ + 2Fe²⁺
   b) Br₂ + 2Fe²⁺ → 2Br⁻ + 2Fe³⁺
7. 3Sn²⁺ + 2NO₃⁻ + 8H⁺ → 2NO + 4H₂O + 3Sn⁴⁺
8. No. The Fe³⁺ would oxidize and dissolve the Al container.
9. a) reverse
   b) forward
10. Mg + 2H⁺ → Mg²⁺ + H₂
11. The H⁺ in acids is not a strong enough oxidizing agent to oxidize silver but the nitrate ion in nitric acid is.
12. 2Na + H₂O → 2Na⁺ + H₂ + OH⁻
   The OH⁻ turns the phenolphthalein pink. The H₂ is ignited by the energy released by the exothermic reaction.
13. a) Ag⁺
   b) Mg
14. Ga³⁺; Al is being oxidized by gallium ions therefore Ga³⁺ is a stronger oxidizing agent (higher on the left in the SRP table) than Al³⁺ and so Ga³⁺ has a greater reduction potential than Al³⁺.
15. a) +1.09V
   b) +0.76V
16. a) Fe + Co²⁺ → Fe²⁺ + Co
   b) Cu⁺ + Cr²⁺ → Cr³⁺ + Cu
   c) Cu²⁺ + Sn²⁺ → Sn⁴⁺ + Cu
17. Zn + 2Fe³⁺ → 2Fe²⁺ + Zn²⁺
18. a) C³⁺
   b) D
19. 2.62 M
20. 0.0179 M
21. a) 2Cr₂O₇²⁻ + C₂H₅OH + 16H⁺ → 4Cr³⁺ + 11H₂O + 2CO₂
   b) 0.0700%. Legally impaired now, not before Sep. 2010.

Page 473, Quick Check
1. Cr
2. from right to left or from the Cr electrode to the Ag electrode
3. from right to left or from the Cr/Cr²⁺ half-cell to the Ag/Ag⁺ half-cell
4. Ag

Page 475, Practice Problems 7.4.1
1. Ni → Ni²⁺ + 2e⁻  \[ E^\circ = 0.26 \text{ V} \]
   Br₂ + 2e⁻ → 2Br⁻  \[ E^\circ = 1.09 \text{ V} \]
   Br₂ + Ni → Ni²⁺ + 2Br⁻  \[ E^\circ = 1.35 \text{ V} \]
2. 2Al → 2Al³⁺ + 6e⁻  \[ E^\circ = 1.66 \text{ V} \]
   3Cu²⁺ + 6e⁻ → 3Cu  \[ E^\circ = 0.34 \text{ V} \]
   2Al + 3Cu²⁺ → 2Al³⁺ + 3Cu  \[ E^\circ = 2.00 \text{ V} \]
3. 6I⁻ → 3I₂ + 6e⁻  \[ E^\circ = -0.54 \text{ V} \]
   2Au³⁺ + 6e⁻ → 2Au  \[ E^\circ = +1.50 \text{ V} \]
   6I⁻ + 2Au³⁺ → 3I₂ + 2Au  \[ E^\circ = +0.96 \text{ V} \]

Page 479, Practice Problems 7.4.2
1. \[ E^\circ_{\text{cell}} = +0.19 \text{ V} \]
   \[ K = 10^{6.4} = 3 \times 10^6 \]
   Prediction: If [Fe³⁺] decreases, Q decreases, log Q decreases, E increases.
   \[ E = +0.23 \text{ V} \]
2. \[ E = 0.52 \text{ V} \]
3. \[ Q = 3.07 \times 10^{11} = 10^{11} \]
   \[ [Cu^{2+}] = \frac{1.0 \text{ M}}{10^{11}} = 10^{-12} \text{ M} \]
**Page 482, Quick Check**

1. hydrated iron (III) oxide or iron(III) oxide monohydrate
2. \[ \frac{1}{2} \text{O}_2(g) + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O} \quad \text{or} \quad \frac{1}{2} \text{O}_2(g) + \text{H}_2\text{O} + 2\text{e}^- \rightarrow 2\text{OH}^- \]
3. impurities or physical stresses or [O₂] or [ions] or surface area
4. coating it with paint or grease or cathodic protection or galvanizing it.

**Page 484, Quick Check**

1. hydrated iron (III) oxide or iron(III) oxide monohydrate
2. \( \frac{1}{2} \text{O}_2(g) + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O} \quad \text{or} \quad \frac{1}{2} \text{O}_2(g) + \text{H}_2\text{O} + 2\text{e}^- \rightarrow 2\text{OH}^- \)
3. impurities or physical stresses or [O₂] or [ions] or surface area
4. coating it with paint or grease or cathodic protection or galvanizing it.

**Page 486, 7.4 Review Questions**

1. | Type            | Anode Material              | Cathode Material            | Electrolyte Medium | Use                                    |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Alkaline Cell</td>
<td>Zn powder packed around a brass pin</td>
<td>MnO₂(s) and powdered C</td>
<td>a moist paste of KOH</td>
<td>e.g. portable electronics, toys, flashlights</td>
</tr>
<tr>
<td>Lead-acid Storage Battery</td>
<td>Pb alloy grids packed with spongy Pb</td>
<td>Pb alloy grids packed with PbO₂</td>
<td>H₂SO₄(aq)</td>
<td>automobiles</td>
</tr>
<tr>
<td>Fuel Cell</td>
<td>porous carbon</td>
<td>porous carbon</td>
<td>proton exchange membrane</td>
<td>e.g. electric vehicles, space travel,</td>
</tr>
</tbody>
</table>

2. a) The diagram should show e⁻ flowing through the wire from the Fe/Fe²⁺ half-cell to the Ni/Ni²⁺ half-cell.
   b) \( \text{Fe/Fe}^{2+} \)
   c) \( \text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^- \); \( \text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni} \)
   d) Fe labelled anode; Ni labelled cathode
   e) \( E^0 = +0.19 \text{ V} \)
   f) \( \text{NH}_4^+ \) flow from the salt bridge into the Ni/Ni²⁺ half-cell
   NO₃⁻ flow through the salt bridge into the Fe/Fe²⁺ half-cell
   g) The Fe electrode loses mass. The Ni electrode gains mass.
3. For example:

4. a) The diagram should show $e^-$ flowing through the wire from the Cr/Cr$^{3+}$ half-cell to the H$_2$/H$^+$ half-cell
b) H$_2$/H$^+$
c) Cr $\rightarrow$ Cr$^{3+}$ + 3$e^-$; 2H$^+$ + 2$e^-$ $\rightarrow$ H$_2$
5. a) Cr labelled anode; Pt labelled cathode
b) 0.74 V
c) Cr$^{3+}$ flow through the porous barrier into the H$_2$/H$^+$ half-cell
Cl$^-$ flow through the porous barrier into the Cr/Cr$^{3+}$ half-cell
d) The Cr electrode loses mass. The Pt electrode’s mass is unchanged.

6.  

7. It allows ion migration that completes the circuit

8. $2Fe^{2+}$ $\rightarrow$ $2Fe^{3+}$ + 2$e^-$  
   $E^\circ$ = –0.77 V  
   $Br_2 + 2e^-$ $\rightarrow$ 2Br$^-$  
   $E^\circ$ = +1.09 V

9. a) 2.07 V (the $E^\circ$ of $F_2 + 2e^-$ $\rightarrow$ 2F$^-$ is 2.07 V greater than the $E^\circ$ of Ag$^+$ + $e^-$ $\rightarrow$ Ag)
b) –3.17 V (the $E^\circ$ of Mg$^{2+}$ + 2$e^-$ $\rightarrow$ Mg is 3.17 V less than the $E^\circ$ of Ag$^+$ + $e^-$ $\rightarrow$ Ag)

10. Cr$^{3+}$ + 2$e^-$ $\rightarrow$ Cr  
    $E^\circ$ = +0.32 V
    Sr $\rightarrow$ Sr$^{2+}$ + 2$e^-$  
    $E^\circ$ = +2.89 V  
    (from Table A7, p.487)
    Cr$^{3+}$ + 2$e^-$ $\rightarrow$ Cr  
    $E^\circ$ = –0.56 V
Sr + Cr^{3+} \rightarrow Sr^{2+} + Cr^{2+} \quad E^\circ = +2.33 \text{ V} \quad (provided by question)

11a) 0.83 V
Cu \rightarrow Cu^{2+} + 2e^- \quad E^\circ = -0.34 \text{ V} \quad (from Table A7, p.487)
Pd^{2+} + 2e^- \rightarrow Pd \quad E^\circ = +0.83 \text{ V}
Cu + Pd^{2+} \rightarrow Cu^{2+} + Pd \quad E^\circ = +0.49 \text{ V} \quad (provided by question)

b) -1.90 V
2Np \rightarrow 2Np^{3+} + 6e^- \quad E^\circ = 1.90 \text{ V}
3Pd^{2+} + 6e^- \rightarrow 3Pd \quad E^\circ = 0.83 \text{ V} \quad (answer to 10 a.)
2Np + 3Pd^{2+} \rightarrow 2Np^{3+} + 3Pd \quad E^\circ = 2.73 \text{ V} \quad (provided by question)

1.90 V is the Standard Oxidation Potential (SOP) of Np therefore the Standard Reduction Potential (SRP) of Np^{3+} is -1.90 V.

12.

<table>
<thead>
<tr>
<th>Reduction Half-cell</th>
<th>Oxidation Half-cell</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cd^{2+} + 2e^- \rightarrow Cd</td>
<td>Pt \rightarrow Pt^{2+} + 2e^-</td>
</tr>
<tr>
<td>Ni^{2+} + 2e^- \rightarrow Ni</td>
<td>Ni \rightarrow Ni^{2+} + 2e^-</td>
</tr>
<tr>
<td>Ce^{3+} + 3e^- \rightarrow Ce</td>
<td>Ce \rightarrow Ce^{3+} + 3e^-</td>
</tr>
</tbody>
</table>

Construct an SRP Table from the data provided:

Pt^{2+} + 2e^- \rightleftharpoons Pt + 1.17 \text{ V} \quad (1.43 \text{ V above Ni}^{2+} + 2e^- \rightarrow \text{Ni})
Ni^{2+} + 2e^- \rightleftharpoons Ni^- + 0.26 \text{ V} \quad (from Table A7, p.487)
Cd^{2+} + 2e^- \rightleftharpoons Cd + 0.43 \text{ V} \quad (0.17 \text{ V below Ni}^{2+} + 2e^- \rightarrow \text{Ni})
Ce^{3+} + 3e^- \rightleftharpoons Ce + 2.36 \text{ V} \quad (1.93 \text{ V below Cd}^{2+} + 2e^- \rightarrow \text{Cd})

13. a) increased voltage (The reactant concentrations are greater than 1 M so the cell’s potential will be greater than the standard cell potential.)
b) 0 V
c) increased voltage (The product concentrations are less than 1 M because the S^{2-} ions precipitate out some Fe^{2+} so the cell’s potential will be greater than the standard cell potential.)
d) no effect

14. For e.g. Reactants biodegradable. Products are just CO_2 and H_2O. No metal or metal ion reactants or products.

15. Weld a zinc bar to the staircase preferably above the area exposed to ocean spray. The zinc is oxidized more readily than the steel (Fe). The zinc bar transfers electrons through the staircase to the oxidizing agents in the ocean spray thus leaving the staircase intact.

16. If the zinc is oxidized it forms zinc oxide. Zinc oxide forms a hard, impenetrable coating. Should the zinc or the zinc oxide coating be scratched off in some areas, the remaining zinc coating still provides cathodic protection.

17. The reactants in electrochemical cells react through electrical contact, via an action-at-a-distance force, without ever coming into physical contact, i.e. colliding.

18. The voltage changes by + 0.03 V each time the concentration changes by a factor of 10.

19. $E_{cell} = -0.86; \quad K = 10^{-1.50}$

20. a) $E_{cell} = +2.20 \text{ V}; \quad K = 10^{2.23}$
b) $E$ is greater in these conditions than standard conditions. Mathematically, Q in standard conditions is 1 (molarities of each ion is 1.0 M) Thus the log of Q is 0. The change of these concentrations resulted in a greater E value. K does not change as there was no change in temperature. (Only changes in temperature can alter K values).
21. a) $E^\circ_{cell} = +0.76 \, V$
   b) $7 \times 10^{-5} \, M$
   c) $pH = 4.2$

22. a) $E^\circ = +0.26 \, V$
   b) $E^\circ_{an} = +0.26 \, V$
   c) Ni(s)

**Page 494, Practice Problems 7.5.1**

1. $\text{2Ag} \rightarrow \text{2Ag}^+ + 2e^- \quad E^\circ = -0.80 \, V$
   $\text{Ni}^{2+} + 2e^- \rightarrow \text{Ni} \quad E^\circ = -0.26 \, V$

   $\text{2Ag} + \text{Ni}^{2+} \rightarrow \text{2Ag}^+ + \text{Ni} \quad E^\circ = -1.06 \, V$

   A voltage of at least 1.06 V would be required to operate this cell.

2. $2F^- \rightarrow F_2 + 2e^- \quad E^\circ = -2.87 \, V$
   $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \quad E^\circ = +0.34 \, V$

   $2F^- + \text{Cu}^{2+} \rightarrow F_2 + \text{Cu} \quad E^\circ = -2.53 \, V$

   A voltage of at least 2.53 V would be required to operate this cell.

3. $3\text{Sn} \rightarrow 3\text{Sn}^{2+} + 6e^- \quad E^\circ = +0.14 \, V$
   $2\text{Al}^{3+} + 6e^- \rightarrow 2\text{Al} \quad E^\circ = -1.66 \, V$

   $3\text{Sn} + 2\text{Al}^{3+} \rightarrow 3\text{Sn}^{2+} + 2\text{Al} \quad E^\circ = -1.52 \, V$

   A voltage of at least 1.52 V would be required to operate this cell.

**Page 496, Quick Check**

1. electrochemical cell
2. electrolytic cell
3. both
4. electrochemical cell
5. both
6. electrolytic cell
7. electrochemical cell

**Page 499, Practice Problems 7.5.2**

1. a) Type 1 $2I^- \rightarrow I_2 + 2e^-$
   $\text{Mg}^{2+} + 2e^- \rightarrow \text{Mg}$
   b) Type 2 $\text{H}_2\text{O} \rightarrow \frac{1}{2} \text{O}_2 + 2\text{H}^+ + 2e^-$
   $2\text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2 + 2 \text{OH}^-$
   c) Type 3 $\text{Fe} \rightarrow \text{Fe}^{3+} + 2e^-$
   $2\text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2 + 2 \text{OH}^-$

2. Anode $2\text{Br}^- \rightarrow \text{Br}_2 + 2e^- \quad E^\circ = -1.09 \, V$
   Cathode $\text{Ni}^{3+} + 2e^- \rightarrow \text{Ni} \quad E^\circ = -0.26 \, V$
   $E^\circ = -1.35 \, V$

A voltage of at least 1.35 V would be required to operate this cell.
Page 501, Quick Check
1. the electrolytic recovery of a metal from a solution containing its ions.
2. aluminum oxide dissolves in molten cryolite (Na3AlF6(\(l\)))
3. sodium hydroxide and chlorine
4. impressed current cathodic protection

Page 503, 7.5 Review Questions
1. citric acid (weak electrolyte), sodium chloride, sodium citrate, and potassium dihydrogen phosphate
2. 

![Diagram of electrolysis cell](image)

Anode \[2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-\] \(E^\circ = -1.36\text{ V}\)
Cathode \[\text{Na}^+ + \text{e}^- \rightarrow \text{Na}\] \(E^\circ = -2.71\text{ V}\)
\[E^\circ = -4.07\text{ V}\]

A voltage of at least 4.07 V would be required to operate this cell.

3. Because the products (Na and Cl\(_2\)) spontaneously react with each other

<table>
<thead>
<tr>
<th>Cell Type (1,2,3)</th>
<th>Electrolyte</th>
<th>Anode/Cathode</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td>Anode</td>
</tr>
<tr>
<td>a.  1</td>
<td>NaCl((l))</td>
<td>Pt / Pt</td>
<td>Cl(_2)</td>
</tr>
<tr>
<td>b.  2</td>
<td>NaCl(aq)</td>
<td>Pt / C</td>
<td>Cl(_2)</td>
</tr>
<tr>
<td>c.  2</td>
<td>CuBr(_2)(aq)</td>
<td>C / C</td>
<td>Br(_2)</td>
</tr>
<tr>
<td>d.  2</td>
<td>AlF(_3)(aq)</td>
<td>C / C</td>
<td>(\frac{1}{2})O(_2) + 2H(^+)</td>
</tr>
<tr>
<td>e.  3</td>
<td>CuCl(_2)(aq)</td>
<td>Cu / Cu</td>
<td>Cu(^{2+})</td>
</tr>
</tbody>
</table>

5. a) Anode \[\text{H}_2\text{O} \rightarrow \frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2\text{e}^-\] \(E^\circ = -1.37\text{ V}\) (approximate)
Cathode \[\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}\] \(E^\circ = +0.96\text{ V}\)
\[E^\circ = -0.41\text{ V}\]

b) A voltage of at least 0.41 V would be required to operate this cell.

6. \[2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-\]

7. In electrorefining, the metallic ions reduced at the cathode are replaced by oxidizing an impure anode.

8. Uneconomical to electrolyze Al\(_2\)O\(_3\)(\(l\)) because aluminum oxide has a high melting point. Impossible to produce Al by electrolyzing Al\(_2\)O\(_3\)(aq) because water is a stronger oxidizing agent than Al\(^{3+}\).

9. In impressed current cathodic protection the metal being protected acts as the cathode in an electrolytic cell (non-spontaneous redox reaction) whereas in galvanic cathodic protection the
metal being protected acts as the cathode in an electrochemical cell (spontaneous redox reaction).

10. 

\[
\begin{align*}
\text{Anode} & \quad \text{Ag} \rightarrow \text{Ag}^+ + e^- \quad \text{E}^\circ = -0.80 \text{ V} \\
\text{Cathode} & \quad \text{Ag}^+ + e^- \rightarrow \text{Ag} \quad \text{E}^\circ = +0.80 \text{ V} \\
\end{align*}
\]

\[\text{E}^\circ = 0.0\]

In theory, it should require no voltage to operate this cell because no net reaction occurs. In practice, a small voltage is required to overcome the internal resistance of the cell.

11. 

a) Type 1 Cell. Only one chemical, water, can be oxidized or reduced. Water is molten H_2O.

b) Any salt containing ions which are weaker reducing agents and weaker oxidizing agents than water; for example, sodium sulphate.

12. 

\[
\begin{align*}
2\text{Br}^- & \rightarrow \text{Br}_2 + 2e^- \quad \text{E}^\circ = -1.09 \text{ V} \\
\text{Cu}^{2+} + 2e^- & \rightarrow \text{Cu} \quad \text{E}^\circ = +0.34 \text{ V} \\
2\text{Br}^- + \text{Cu}^{2+} & \rightarrow \text{Br}_2 + \text{Cu} \quad \text{E}^\circ = -0.75 \text{ V} \\
\end{align*}
\]

Page 506, Practice Problems

1. 121 g Al (s)
2. 1.5 A
3. a) This reaction takes place at the anode because species lose electrons at the anode resulting in oxidation. Sulfate is losing electrons in this half reaction.
   b) 3.15 h

Page 509, Practice Problems

1. \(\Delta G^\circ = -46 \text{ kJ/mol}_{\text{rxn}}\)
2. \(-0.40 \text{ V}\)
3. a) \(-2.19 \text{ V}\)
   b) \(E^-_{\text{ox}} = -0.53 \text{ V}\)
   \(X_2 = I_2 = \) diatomic iodine
   If \(-0.53 \text{ V}\) is the oxidation potential, the reduction potential is \(+0.53 \text{ V}\). The oxidation potential corresponds to the iodide ion.

Page 511, Review Questions

1. a) 0.44 g Cr (s)
   b) 820 s or 7.3 h
   c) 1.99 A

2. a) Placing a layer of tin against the steel, which is mainly composed of iron, protects the iron/steel can. Iron oxidizes spontaneously with water (and oxygen) where tin will not spontaneously react, thus becoming a protective barrier. Furthermore, if the iron were to oxidize at all, it would force electrons onto the tin keeping the tin in its reduced form. The first “layer” of protection is “barrier protection”. The second is called “cathodic protection”.
   b) i) 6.92 g Sn (s)
   ii) 3.69 M
3. 2080 s
4. a) 5.38 g Co (s)
   b) 1.02 L O² (g)
   c) Two methods to identify that the gas is oxygen gas:
      1 - relighting a glowing splint
      2 - using an indicator to show the accompanying H⁺ ions
5. a) $E^{\circ}_{\text{cell}} = -1.23 \text{ V}$, $\Delta G^{\circ} = 475 \text{ kJ/mol}_{\text{rxn}}$
   b) $E^{\circ}_{\text{cell}} = +1.27 \text{ V}$, $\Delta G^{\circ} = -490 \text{ kJ/mol}_{\text{rxn}}$. (b) is spontaneous
6. a) $E^{\circ}_{\text{red}} = +0.34 \text{ V}$, $E^{\circ}_{\text{red}} = +0.34 \text{ V}$
   b) $\Delta G^{\circ} = -213 \text{ kJ/mol}_{\text{rxn}}$
7. a) $\Delta G^{\circ} = -89 \text{ kJ/mol}_{\text{rxn}}$
   b) $-0.17 \text{ V}$
8. 4.3 g
9. $E^{\circ} = +1.33 \text{ V}$
10. 9.30 V
11. a) $6 \text{ AlO}_3^{2-} + 3 \text{ C(s)} \rightarrow 2 \text{ AlF}_6^{3-} + 3 \text{ CO}_2 + 4 \text{ Al}$
   b) $\Delta G = 320 \text{ kJ/mol}_{\text{rxn}}$