Chapter 7 - Answer Key

Page 433, Practice Problems

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"?
   my partner (they got reduced)
   A "reducing agent"?   me (I got oxidized)
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

0   0   -4 +1
1. $C(s) + 2H_2(g) \rightarrow CH_4(g)$ O.A. = $C(s)$
   R.A. = $H_2(g)$

0   +3 -1   0   +2 -1
2. $3Sr(s) + 2FeBr_3(aq) \rightarrow 2Fe(s) + 3SrBr_2(aq)$
   O.A. = Fe^{3+} R.A. = Sr

+2 -2   +5 -2   +4 -2   0
3. $5CO(g) + Cl_2O_5(s) \rightarrow 5CO_2(g) + Cl_2(s)$
   O.A. = Cl_2O_5 R.A. = CO

-3 +1   0   0
4. $4PH_3(g) \rightarrow P_4(g) + 6H_2(g)$
   O.A. and R.A. = PH_3

0   +1 -2   +2 -2 +1   0
5. $Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(s) + H_2(g)$
   O.A. = H_2O R.A. = Ba

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) $\underline{Ca}I_2$ b) $\underline{O}F_2$ c) $\underline{C}_6H_{12}O_6$ d) $Rb_2\underline{O_2}$
   +2   +2   0   -1
   e) $\underline{S}_2O_3^{2-}$ f) $Be\underline{H}_2$ g) $\underline{Br}O^-$ h) $\underline{Cl_2}$
   +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.

+1 +5 -2   +1 -1   0
4. a) $2KBrO_3(s) \rightarrow 2KBr(s) + 3O_2(g)$
   Oxidized: $KBrO_3$ Reduced: $KBrO_3$

0   +1 +5 -2   +2 +5 -2   0
b) $Sr(s) + 2CuNO_3(aq) \rightarrow Sr(NO_3)_2(aq) + 2Cu(s)$
   Ox'd: Sr   Red'd: Cu^+

0   0   +2 -1
c) $2F_2(g) + O_2(g) \rightarrow 2OF_2(g)$
   Oxidized: O_2 Reduced: F_2

-3 +1+5 -2   +1 -2   +1 -2
d) $NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l)$
   Ox'd: NH_4NO_3 Red'd: NH_4NO_3 (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) $2Sn(s) + O_2(g) \rightarrow 2SnO(s)$
OA: O_2 RA: Sn   No e^-: 2

0   0   +5 -1
b) $2V(s) + 5I_2(g) \rightarrow 2VI_5(s)$
OA: I_2 RA: V   No e^-: 5

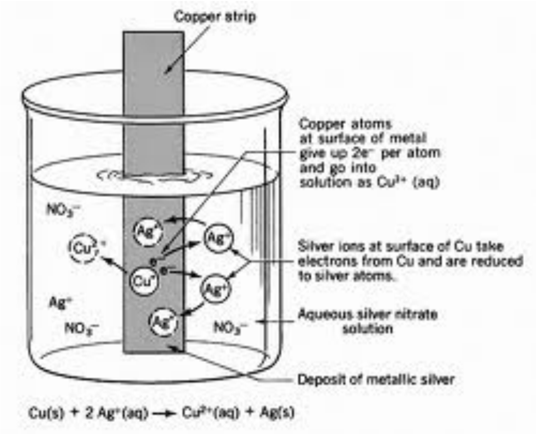
0   +1 -1   +2 -1   0
c) $Sr(s) + 2HCl(aq) \rightarrow SrCl_2(aq) + H_2(g)$
OA: H+   RA: Sr   No e^-: 2

-8/3 +1   0   +4 -2   +1 -2
d) $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
OA: O_2 RA: C_3H_8 No e^-: 6⅔

6. a) Write a balanced redox equation (in net ionic form) to show what has occurred in the beaker over time: Fe(s) + Cu2+(aq) → Fe2+(aq) + Cu(s)
   b) What is the oxidizing agent? Cu2+(aq)
   The reducing agent? 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) $\underline{P}^{3-}$ -3 b) $(NH_4)_2\underline{Zr}(SO_4)_3$ +4 c) $Na_2\underline{C}_2O_4$ +3
d) $\underline{N}_2H_5Cl$ -2 e) $\underline{Mn}O_4^{2-}$ +6

8.



a) see above
b) Oxidized? Cu metal
Reduced? Ag^+ ion
c) Reducing agent? Cu metal
Oxidizing agent? Ag^+ ion
d) 2e-

9. a) Alakli metal family (IA)
b) Halogents (VIIA) or (17)

10. a) ClO_4^-, ClO_3^-, ClO_2^-
b) Cl_2, Cl^-

Page 444, Practice Problems

1. $Sm \rightarrow Sm^{3+} + 3e^-$ *oxidation*
2. $8e^- + 10\,H^+ + NO_3^- \rightarrow NH_4^+ + 3\,H_2O$ *reduction*
3. $2e^- + H_2O + IO_4^- \rightarrow IO_3^- + 2OH^-$ *reduction*
4. $5\,H_2O + S_2O_3^{2-} \rightarrow 2\,SO_4^{2-} + 10\,H^+ + 8e^-$ *oxidation*
5. $10e^- + 6\,H_2O + 2\,BrO_3^- \rightarrow Br_2 + 12\,OH^-$ *reduction*

Page 447, Practice Problems

1. $4\,Zn + 8\,H^+ + H_3AsO_4 \rightarrow AsH_3 + 4\,H_2O + 4\,Zn^{2+}$

Species	**Reactants**	**Products**
Hydrogen	11	11
Zinc	4	4
Oxygen	4	4
Arsenic	1	1
Charge	+8	+8

2.
$4\,H^+ + 4\,NO_3^- + C_2H_5OH \rightarrow CH_3COOH + 2\,N_2O_4 + 3\,H_2O$

Species	**Reactants**	**Products**
Hydrogen	10	10
Nitrogen	4	4
Oxygen	13	13
Carbon	2	2
Charge	0	0

Page 449, Practice Problems 7.2.3

1. $2\,HXeO_4^- + 2\,OH^- \rightarrow XeO_6^{4-} + Xe + O_2 + 2\,H_2O$

Species	**Reactants**	**Products**
Hydrogen	4	4
Xenon	2	2
Oxygen	10	10
Charge	-4	-4

2. $6\,H_2O + 2\,BrO_3^- + 10\,Br^- \rightarrow 6\,Br_2 + 12\,OH^-$

Species	**Reactants**	**Products**
Hydrogen	12	12
Bromine	12	12
Oxygen	12	12
Charge	-12	-12

3. $2\,H_2O + 2\,CH_3COO^- \rightarrow 2\,CH_4 + 2\,CO_2 + 2\,OH^-$

Species	**Reactants**	**Products**
Hydrogen	10	10
Carbon	4	4
Oxygen	6	6
Charge	-2	-2

Page 452, 7.2 Review Questions

1.

	Reactants	Products	Add
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

			products
d.	1+	5+	4 e^- to products

2\.

a) $2 e^- + 2 NO_3^- + 2 H_2O \rightarrow N_2O_4 + 4 OH^-$ (reduction)

b) $2 Cr^{3+} + 7 H_2O \rightarrow Cr_2O_7^{2-} + 14 H^+ + 6 e^-$ (oxidation)

c) $8 e^- + ClO_4^- + 4 H_2O \rightarrow Cl^- + 8 OH^-$ (reduction)

d) $S_2O_5^{2-} + 3 H_2O \rightarrow 2 SO_4^{2-} + 6 H^+ + 4 e^-$ (oxidation)

3\. a) $14 e^- + 16 H^+ + 2 ClO_4^- \rightarrow Cl_2 + 8 H_2O$ (reduction)

b) $H_2O + 2 FeO \rightarrow Fe_2O_3 + 2 H^+ + 2 e^-$ (oxidation)

c) $2 H_2O + N_2O_4 \rightarrow 2 NO_3^- + 4 H^+ + 2 e^-$ (oxidation)

4\. a) $4 H_2O + 4 e^- + CrO_4^{2-} \rightarrow Cr(OH)_2 + 6 OH^-$ (reduction)

b) $2 S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2 e^-$ (oxidation)

c) $4 e^- + 3 H_2O + IO_3^- + 2 Cl^- \rightarrow ICl_2^- + 6 OH^-$ (reduction)

5\. a) $4 H_2O + 2 MnO_4^- + 3 Sn^{2+} \rightarrow 3 Sn^{4+} + 2 MnO_2 + 8 OH^-$

b) $2 H^+ + 2 H_2SO_3 + 4 V^{2+} \rightarrow 4 V^{3+} + S_2O_3^{2-} + 3 H_2O$

c) $3 H_2O + IO_3^- + 5 I^- \rightarrow 3 I_2 + 6 OH^-$

d) $3 N_2H_4 + 4 ClO_3^- \rightarrow 4 Cl^- + 6 NO + 6 H_2O$

e) $4H_2O + 2NO_3^- + 3Zn \rightarrow 3Zn^{2+} + 2NO + 8OH^-$

f) $2 H^+ + 11 ClO_3^- \rightarrow 8 ClO_4^- + Cl^- + Cl_2 + H_2O$

g) $5 H_2O + 3 SnS_3O_3 + 16 MnO_4^- \rightarrow 16 MnO_2 + 10 OH^- + 9 SO_4^{2-} + 3 Sn^{4+}$

h) $2 Mg_3(AsO_4)_2 + 6 SiO_2 + 10 C \rightarrow 10 CO + As_4 + 6 MgSiO_3$

6\. a) $6H^+ + SeO_3^{2-} + 4F^- \rightarrow Se + 2F_2 + 3H_2O$

b) $2 H_2O + 4 ReO_4^- + 3 Sb_2O_3 \rightarrow 4 ReO_2 + 3 Sb_2O_5 + 4 OH^-$

c) $16 H^+ + 4 NO_3^- + 3 Pd + 18 I^- \rightarrow 3 PdI_6^{2-} + 4 NO + 8 H_2O$

d) $Pb(OH)_4^{2-} + BrO^- \rightarrow PbO_2 + Br^- + H_2O + 2 OH^-$

7\. a) $16 HCl + 2 K_2Cr_2O_7 + 3 CH_3CH_2OH \rightarrow$
$3 CH_3COOH + 4 CrCl_3 + 11 H_2O + 4 KCl$

b) $4 NaCl + 2H_2SO_4 + MnO_2 \rightarrow$
$2 Na_2SO_4 + MnCl_2 + 2 H_2O + Cl_2$

Page 459, Practice Problems

1\. yes; $2I^- + Br_2 \longrightarrow 2Br^- + I_2$

2\. no; both are oxidizing agents

3\. yes; $2Ag^+ + Sn \longrightarrow Sn^{2+} + 2Ag$

4\. no; I_2 isn't a strong enough oxidizing agent to spontaneously react with Cl^-

Page 460, Practice Problems

1\. no; I_2 isn't a strong enough oxidizing agent to spontaneously react with F^-

2\. yes; $3Cu^{2+} + 2Al \longrightarrow 2Al^{3+} + 3Cu$

3\. yes; $Cl_2 + 2Br^- \longrightarrow Br_2 + 2Cl^-$ (* 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

1\. $Sn^{4+} + Ni \longrightarrow Sn^{2+} + Ni^{2+}$

2\. $3Cu^{2+} + 2Al \longrightarrow 2Al^{3+} + 3Cu$

3\. $2Cu^+ \longrightarrow Cu^{2+} + Cu$

4\. $Cu + Br_2 \longrightarrow 2Br^- + Cu^{2+}$

Page 465, Practice Problems

1\. a) $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

b) 0.0900 M

2\. 0.183 M

3\. a) 1.97 M

b) 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; $3Mg + 2Al^{3+} \longrightarrow 3Mg^{2+} + 2Al$

b) no

c) yes; $Hg^{2+} + 2Ag \longrightarrow Hg + 2Ag^+$

4.

Metals	**Non-metals**
bottom right of SRP Table	**top left of SRP Table**
tend to give electrons	tend to take electrons
give e^- to chemicals above them on the left	take e^- from chemicals below them on the right

5. a) yes; $Fe + Sn^{2+} \rightarrow Fe^{2+} + Sn$
   b) yes; $F_2 + 2Br^- \rightarrow Br_2 + 2F^-$
   c) no
6. a) $2I^- + 2Fe^{3+} \rightarrow I_2 + 2Fe^{2+}$
   b) $Br_2 + 2Fe^{2+} \rightarrow 2Br^- + 2Fe^{3+}$
7. $3Sn^{2+} + 2NO_3^- + 8H^+ \rightarrow 2NO + 4H_2O + 3Sn^{4+}$
8. No. The Fe^{3+} would oxidize and dissolve the Al container.
9. a) reverse
   b) forward
10. $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$
11. The H^+ in acids is not a strong enough oxidizing agent to oxidize silver but the nitrate ion in nitric acid is.
12. $2Na + H_2O \rightarrow 2Na^+ + H_2 + OH^-$
    The OH^- turns the phenolphthalein pink. The H_2 is ignited by the energy released by the exothermic reaction.
13. a) Ag^+
    b) Mg
14. Ga^{3+}; Al is being oxidized by gallium ions therefore Ga^{3+} is a stronger oxidizing agent (higher on the left in the SRP table) than Al^{3+} and so Ga^{3+} has a greater reduction potential than Al^{3+}.
15. a) +1.09V
    b) +0.76V
16. a) $Fe + Co^{2+} \rightarrow Fe^{2+} + Co$
    b) $Cu^+ + Cr^{2+} \rightarrow Cr^{3+} + Cu$
    c) $Cu^{2+} + Sn^{2+} \rightarrow Sn^{4+} + Cu$
17. $Zn + 2Fe^{3+} \rightarrow 2Fe^{2+} + Zn^{2+}$
18. a) C^{3+}
    b) D
19. 2.62 M
20. 0.0179 M
21. a) $2Cr_2O_7^{2-} + C_2H_5OH + 16H^+ \rightarrow 4Cr^{3+} + 11H_2O + 2CO_2$
    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/Cr^{3+} half-cell to the Ag/Ag^+ half-cell
4. Ag

Page 475, Practice Problems

1. $Ni \rightarrow Ni^{2+} + 2e^-$ E° = 0.26 V
   $Br_2 + 2e^- \rightarrow 2Br^-$ E° = 1.09 V
   $Br_2 + Ni \rightarrow Ni^{2+} + 2Br^-$ E° = 1.35 V

2. $2Al \rightarrow 2Al^{3+} + 6e^-$ E° = 1.66 V
   $3Cu^{2+} + 6e^- \rightarrow 3Cu$ E° = 0.34 V
   $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$ E° = 2.00 V

3. $6I^- \rightarrow 3I_2 + 6e^-$ E° = −0.54 V
   $2Au^{3+} + 6e^- \rightarrow 2Au$ E° = +1.50 V
   $6I^- + 2Au^{3+} \rightarrow 3I_2 + 2Au$ E° = +0.96 V

Page 479, Practice Problems

1. $E^\circ_{cell} = +0.19$ V;
   $K = 10^{6.4} = 3 \times 10^6$
   Prediction: If $[Fe^{2+}]$ decreases, Q decreases, log Q decreases, E increases. $E = +0.23\ V$
2. $E = 0.52\ V$
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.

Type	Anode Material	Cathode Material	Electrolyte Medium	Use
Alkaline Cell	Zn powder packed around a brass pin	$MnO_2(s)$ and powdered C	a moist paste of KOH	e.g. portable electronics, toys, flashlights
Lead-acid Storage Battery	Pb alloy grids packed with spongy Pb	Pb alloy grids packed with PbO_2	$H_2SO_4(aq)$	automobiles
Fuel Cell	porous carbon	porous carbon	proton exchange membrane	e.g. electric vehicles, space travel,

Page 484, Quick Check

1. hydrated iron (III) oxide or iron(III) oxide monohydrate
2. $\frac{1}{2}O_2(g) + 2H^+ + 2e^- \rightarrow H_2O$ or $\frac{1}{2}O_2(g) + H_2O + 2e^- \rightarrow 2OH^-$
3. impurities or physical stresses or $[O_2]$ or [ions] or surface area
4. coating it with paint or grease or cathodic protection or galvanizing it.

Page 486, 7.4 Review Questions

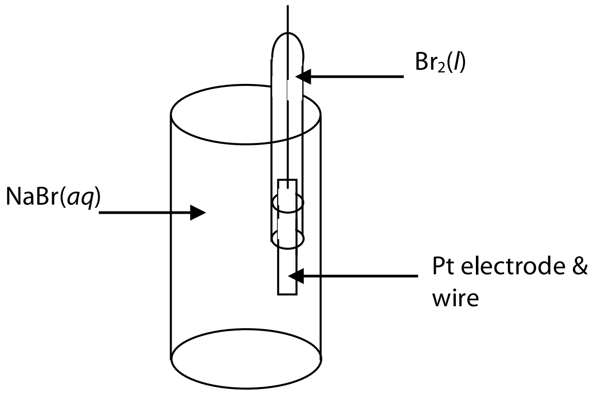
1.

Anode	Cathode
oxidation occurs	reduction occurs
mass decreases	**mass increases**
attracts anions	attracts cations
electrons flow away	**electrons flow towards**

2. a) The diagram should show e^- flowing through the wire from the Fe/Fe^{2+} half-cell to the Ni/Ni^{2+} half-cell.
b) Fe/Fe^{2+}
c) $Fe \rightarrow Fe^{2+} + 2e^-$; $Ni^{2+} + 2e^- \rightarrow Ni$
d) Fe labelled anode; Ni labelled cathode
e) E° = +0.19 V

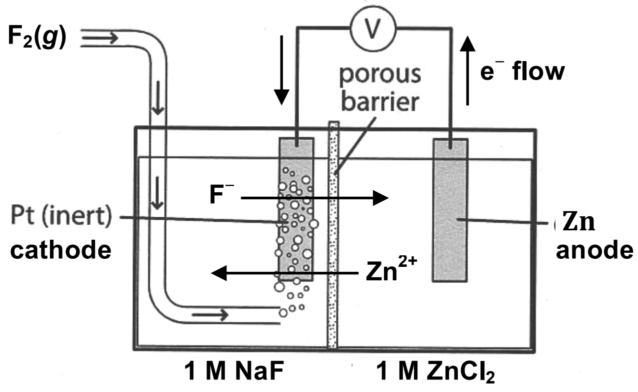
f) NH_4^+ flow from the salt bridge into the Ni/Ni^{2+} half-cell
NO_3^- flow through the salt bridge into the Fe/Fe^{2+} half-cell
g) The Fe electrode loses mass. The Ni electrode gains mass.

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) $Cr \rightarrow Cr^{3+} + 3e^-$; $2H^+ + 2e^- \rightarrow H_2$
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.



d) $Zn \rightarrow Zn^{2+} + 2e^-$ E° = 0.76 V
$F_2 + 2e^- \rightarrow 2F^-$ E° = 2.87 V

e) $Zn + F_2 \rightarrow Zn^{2+} + 2F^-$ E° = 3.63 V

7. It allows ion migration that completes the circuit

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

$2Fe^{2+} + Br_2 \rightarrow 2Fe^{3+} + 2Br^-$ E° = +0.32 V

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

10. $Cr^{2+} + 2e^- \rightarrow Cr$ E° = −0.56 V
$Sr \rightarrow Sr^{2+} + 2e^-$ E° = +2.89 V (from Table A7, p.487)
$Cr^{2+} + 2e^- \rightarrow Cr$ E° = −0.56 V

$Sr + Cr^{2+} \longrightarrow Sr^{2+} + Cr$ E° = +2.33 V (provided by question)

11a) 0.83 V

$Cu \longrightarrow Cu^{2+} + 2e^{-}$ E° = −0.34 V (from Table A7, p.487)
$Pd^{2+} + 2e^{-} \longrightarrow Pd$ E° = +0.83 V

$Cu + Pd^{2+} \longrightarrow Cu^{2+} + Pd$ E° = +0.49 V (provided by question)

b) −1.90 V

$2Np \longrightarrow 2Np^{3+} + 6e^{-}$ E° = 1.90 V
$3Pd^{2+} + 6e^{-} \longrightarrow 3Pd$ E° = 0.83 V (answer to 10 a.)

$2Np + 3Pd^{2+} \longrightarrow 2Np^{3+} + 3Pd$ E° = 2.73 V (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.

		Reduction Half-cell	
		$Cd^{2+} + 2e^{-} \longrightarrow Cd$	$Pt^{2+} + 2e^{-} \longrightarrow Pt$
Oxidation Half-cell	$Pt \longrightarrow Pt^{2+} + 2e^{-}$	**– 1.60 V**	0 V
	$Ni \longrightarrow Ni^{2+} + 2e^{-}$	– 0.17 V	+ 1.43 V
	$Ce \longrightarrow Ce^{3+} + 3e^{-}$	+ 1.93 V	**+ 3.53 V**

Construct an SRP Table from the data provided:

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

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. $E^{\circ}_{cell} = -0.86$; $K = 10^{-150}$
20. a) $E^{\circ}_{cell} = +2.20$ V; $K = 10^{223}$

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) $E = +0.26\ V$
    b) $E^\circ_{ox} = +0.26$ V
    c) Ni(s)

Page 494, Practice Problems

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

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

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

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

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

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

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

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

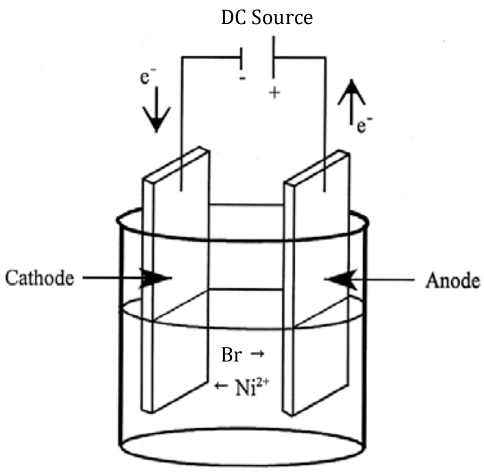
Page 499, Practice Problems

1. a) Type 1 $2\ I^- \rightarrow I_2 + 2e^-$
$Mg^{2+} + 2e^- \rightarrow Mg$
b) Type 2 $H_2O \rightarrow \frac{1}{2}\ O_2 + 2H^+ + 2e^-$
$2H_2O + 2e^- \rightarrow H_2 + 2\ OH^-$
c) Type 3 $Fe \rightarrow Fe^{2+} + 2e^-$
$2H_2O + 2e^- \rightarrow H_2 + 2\ OH^-$

2. Anode $2Br^- \rightarrow Br_2 + 2e^-$ $E^\circ = -1.09$ V
Cathode $Ni^{2+} + 2e^- \rightarrow Ni$ $E^\circ = -0.26$ V

$E^\circ = -1.35$ V

A voltage of at least 1.35 V would be required to operate this cell.

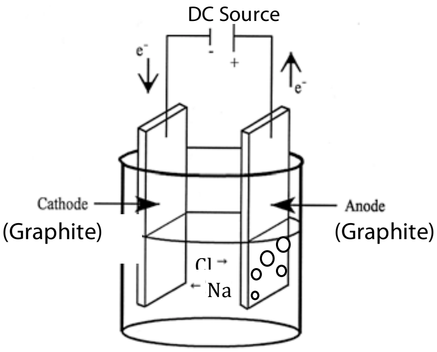


Page 501, Quick Check

1. the electrolytic recovery of a metal from a solution containing its ions.
2. aluminum oxide dissolves in molten cryolite ($Na_3AlF_6(l)$)
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.



Anode $2Cl^- \rightarrow Cl_2 + 2e^-$ $E° = -1.36\ V$

Cathode $Na^+ + e^- \rightarrow Na$ $E° = -2.71\ V$

$E° = -4.07\ V$

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
4.

	Cell Type (1,2,3)	Electrolyte	Anode/Cathode	Products	
				Anode	Cathode
a.	**1**	$NaCl(l)$	Pt / Pt	**Cl_2**	**Na**
b.	**2**	$NaCl(aq)$	Pt / C	**Cl_2**	**$H_2 + 2OH^-$**
c.	**2**	$CuBr_2(aq)$	C / C	**Br_2**	**Cu**
d.	**2**	$AlF_3(aq)$	C / C	**$\frac{1}{2}O_2 + 2H^+$**	**$H_2 + 2OH^-$**
e.	**3**	$CuCl_2(aq)$	Cu / Cu	**Cu^{2+}**	**Cu**

5. a) Anode $H_2O \rightarrow \frac{1}{2}\,O_2 + 2H^+ + 2e^-$ $E° = -1.37\ V$ (approximate)

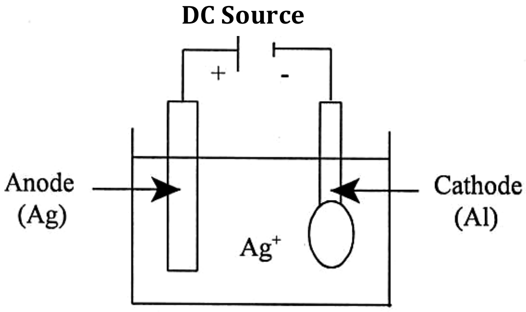
   Cathode $NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ $E° = +0.96\ V$

   $E° = -0.41\ V$

   b) A voltage of at least 0.41 V would be required to operate this cell.
6. $2H_2O + 2e^- \rightarrow H_2 + 2\,OH^-$
7. In electrorefining, the metallic ions reduced at the cathode are replaced by oxidizing an impure anode.
8. Uneconomical to electrolyze $Al_2O_3(l)$ because aluminum oxide has a high melting point. Impossible to produce Al by electrolyzing $Al_2O_3(aq)$ 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.



$Ag \rightarrow Ag^+ + e^-$ $E° = -0.80$ V

$Ag^+ + e^- \rightarrow Ag$ $E° = +0.80$ V

$E° = 0.0$

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. 0.75 V

$2Br^- \rightarrow Br_2 + 2e^-$ $E° = -1.09$ V

$Cu^{2+} + 2e^- \rightarrow Cu$ $E° = +0.34$ V

$2Br^- + Cu^{2+} \rightarrow Br_2 + Cu$ $E° = -0.75$ V

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. $\Delta G° = -46\ kJ/mol_{rxn}$
2. – 0.40 V
3. a) – 2.19 V
   b) $E°_{ox} = -0.53$ V
   $X_2 = I_2 =$ diatomic Iodine
   If – 0.53 V is the oxidation potential, the reduction potential is + 0.53 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 O_2 (g)
   c) Two methods to identify hat the gas is oxygen gas: 1 - relighting a glowing splint
   2 - using an indicator to show the accompanying H^+ ions
5. a) $E^\circ_{cell} = -1.23$ V, $\Delta G^\circ = 475\ kJ/mol_{rxn}$
   b) $E^\circ_{cell} = 1.27$ V, $\Delta G^\circ = -490.\ kJ/mol_{rxn}$. (b) is spontaneous
6. a) $E^\circ_{red} = +0.34$ V, $E^\circ_{red} = +0.34$ V
   b) $\Delta G^\circ = -213\ kJ/mol_{rxn}$
7. a) $\Delta G^\circ = -/\ 80\ kJ/mol_{rxn}$
   b) – 0.17 V
8. 4.3 g
9. $E^\circ = +1.33\ V$
10. 9.30 V
11. a) $6\ AlOF_3^{2-} + 3\ C(s) \rightarrow 2\ AlF_6^{3-} + 3\ CO_2 + 4\ Al$
    b) $\Delta G = 5240\ kJ/mol_{rxn}$