BC Science Chemistry 12

Chapter 6 – Oxidation – Reduction and Its Applications Answer Key

September 20, 2012

6.1 Oxidation – Reduction

Warm Up

Name	Formula	Charge on Metal Ion
iron(III) chloride	$FeCl_3$	**+3**
iron(II) acetate	$Fe(CH_3COO)_2$	**+2**
manganese(II) bromate	$Mn(BrO_3)_2$	**+2**
manganese(IV) oxide	MnO_2	**+4**
potassium permanganate	$KMnO_4$	**+7**

Practice Problems:

1. H_2O H = **+1** O = **-2**
2. Cs_2O_2 Cs = **+1** O = **-1**
3. CO_3^{2-} C = **+4** O = **-2**
4. $Na_2Cr_2O_7$ Na = **+1** Cr = **+6** O = **-2**
5. BaH_2 Ba = **+2** H = **-1**
6. NH_4^+ N = **-3** H = **+1**
7. S_8 S = **0**
8. $Al_2(SO_4)_3$ Al = **+3** S = **+6** O = **-2**

Quick Check:

1. **6** e-
2. **nothing (no e- transferred)**

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)**

Practice Problems: Agents

1. $\overset{0}{C}(s) + 2\,\overset{0}{H_2}(g) \rightarrow \overset{-4}{C}\overset{+1}{H_4}(g)$ O.A. = $C(s)$ R.A. = $H_2(g)$
2. $3\,\overset{0}{Sr}(s) + 2\,\overset{+3}{Fe}\overset{-1}{Br_3}(aq) \rightarrow 2\,\overset{0}{Fe}(s) + 3\,\overset{+2}{Sr}\overset{-1}{Br_2}(aq)$ O.A. = Fe^{3+} R.A. = Sr
3. $5\,\overset{+2}{C}\overset{-2}{O}(g) + \overset{+5}{Cl_2}\overset{-2}{O_5}(s) \rightarrow 5\,\overset{+4}{C}\overset{-2}{O_2}(g) + \overset{0}{Cl_2}(s)$ O.A. = Cl_2O_5 R.A. = CO
4. $4\,\overset{-3}{P}\overset{+1}{H_3}(g) \rightarrow \overset{0}{P_4}(g) + 6\,\overset{0}{H_2}(g)$ O.A. and R.A. = PH_3
5. $\overset{0}{Ba}(s) + 2\,\overset{+1}{H_2}\overset{-2}{O}(l) \rightarrow \overset{+2}{Ba}(\overset{-2}{O}\overset{+1}{H})_2(s) + \overset{0}{H_2}(g)$ O.A. = H_2O R.A. = Ba

6.1 Activity

Characteristic	**Oxidizing Agent**	**Reducing Agent**
Causes another species to be…	oxidized	reduced
Is itself _____ during reaction.	reduced	oxidized
Its oxidation state is__during reaction.	decreased	increased
Causes electrons to be...	lost	gained
_____ electrons.	gains	loses
Belongs to _____ family.	halogen	alkali METAL
Has a _____electronegativity.	high	low

6.1 Review Questions

1. 1. Elements that get oxidized (act as reducing agents) form_(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. $\underline{Ca}I_2$ b) $\underline{O}F_2$ c) $\underline{C}_6H_{12}O_6$ d) $Rb_2\underline{O}_2$ e) $\underline{S}_2O_3^{2-}$ f) $Be\underline{H}_2$ g) $\underline{Br}O^-$ h) $\underline{Cl}_2$

   +2 +2 0 -1 +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 meakes 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.

a) $2\ \overset{+1\ +5\ -2}{KBrO_3}(s) \rightarrow 2\ \overset{+1\ -1}{KBr}(s) + 3\ \overset{0}{O_2}(g)$ Oxidized: **$KBrO_3$** Reduced: **$KBrO_3$**

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

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

d) $\overset{-3\ +1+5\ -2}{NH_4NO_3}(s) \rightarrow \overset{+1\ -2}{N_2O}(g) + 2\ \overset{+1\ -2}{H_2O}(l)$ 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.

(a) $2\ \overset{0}{Sn}(s) + \overset{0}{O_2}(g) \rightarrow 2\ \overset{+2\ -2}{SnO}(s)$ OA: **O_2** RA: **Sn** No e^-: **2**

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

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

(d) $\overset{-8/3\ +1}{C_3H_8}(g) + 5\ \overset{0}{O_2}(g) \rightarrow 3\ \overset{+4\ -2}{CO_2}(g) + 4\ \overset{+1\ -2}{H_2O}(g)$ 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: **$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$**

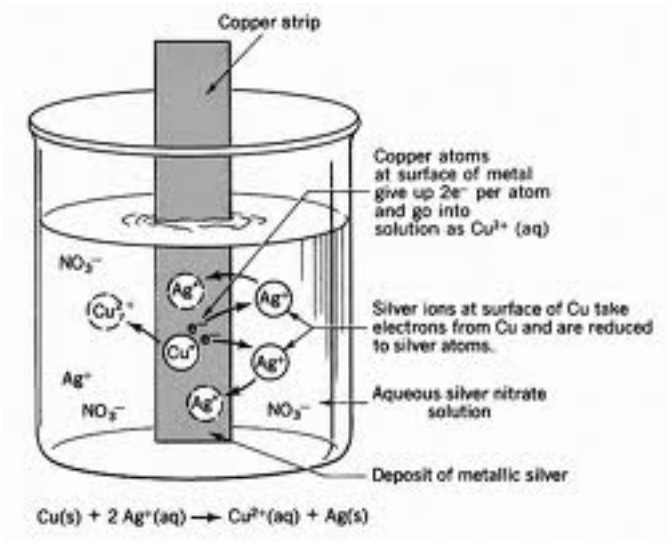
b) What is the oxidizing agent? **$Cu^{2+}(aq)$** The reducing agent? **$Fe(s)$**

c) How many electrons were transferred in the equation? **2**e-

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) $Na_2\underline{C_2}O_4$ **+3** d) $\underline{N_2}H_5Cl$ **-2** e) $\underline{Mn}O_4^{2-}$ **+6**

8.



a) Write a balanced equation to represent the redox reaction occurring in the beaker. (see above)

b) Oxidized? **Cu metal** Reduced? **Ag^+ ion**

c) Reducing agent? **Cu metal** Oxidizing agent? **Ag^+ ion**

d) **2**e-

a) **Alkali metal family (IA)**

b) **Halogens (VIIA) or (17)**

9. a) **ClO_4^-, ClO_3^-, ClO_2^-**

b) **Cl_2, Cl^-**

6.2 Balancing Oxidation-Reduction Equations

Warm Up

Does this equation appear to be balanced? **_yes or no_**

Which species is oxidized and consequently acts as a reducing agent? **__Zn___**

Which species is reduced and consequently acts as an oxidizing agent? **__Cu^+___**

How many electrons are lost by each atom of reducing agent being oxidized? **_2_**

How many electrons are gained by each ion of the oxidizing agent as it is reduced? **_1_**

Rewrite the equation in its balanced form: **$Zn(s) + 2\ Cu^+(aq) \rightarrow Zn^{2+}(aq) + 2\ Cu(s)$**

Practice Problems:

1. **$Sm \rightarrow Sm^{3+} + 3e^-$** ***oxidation***

2. **$8e^- + 10\ H^+ + NO_3^- \rightarrow NH_4^+ + 3\ H_2O$** ***reduction***

3. $2e^- + \cancel{2\ H^+} + IO_4^- \rightarrow IO_3^- + \cancel{H_2O}$
   $\cancel{2\ OH^-}$ $2OH^-$

   $2e^- + H_2O + IO_4^- \rightarrow IO_3^- + 2OH^-$ ***reduction***

4. **$5\ H_2O + S_2O_3^{2-} \rightarrow 2\ SO_4^{2-} + 10\ H^+ + 8e^-$** ***oxidation***

5. $10e^- + 12\ H^+ + 2\ BrO_3^- \rightarrow Br_2 + \cancel{6\ H_2O}$
   $12\ OH^-$ $12\ OH^-$

   $10e^- + 6\ H_2O + 2\ BrO_3^- \rightarrow Br_2 + 12\ OH^-$ ***reduction***

Practice Problems

1. $8e^- + 8\ H^+ + H_3AsO_4 \rightarrow AsH_3 + 4\ H_2O$
   $(Zn \rightarrow Zn^{2+} + 2e^-)\ \times 4$

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

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

2. $H_2O + C_2H_5OH \rightarrow CH_3COOH + 4\ H^+ + 4e^-$
   $(2e^- + 4\ H^+ + 2\ NO_3^- \rightarrow N_2O_4 + 2\ H_2O)\ \times 2$

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

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

Practice Problems – Disproportionation

1. $2\ H_2O + HXeO_4^- \rightarrow XeO_6^{4-} + 5\ H^+ + 2e^-$
   $2e^- + 3\ H^+ + HXeO_4^- \rightarrow Xe + O_2 + 2\ H_2O$

   $2\ HXeO_4^- \rightarrow XeO_6^{4-} + Xe + O_2 + 2\ H^+$
   $2\ OH^-$ $\qquad$ $2\ OH^-$

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

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

2. $10e^- + 12\ H^+ + 2\ BrO_3^- \rightarrow Br_2 + 6\ H_2O$
   $(2\ Br^- \rightarrow Br_2 + 2e^-)$ x 5

   $12\ H^+ + 2\ BrO_3^- + 10\ Br^- \rightarrow 6\ Br_2 + 6\ H_2O$
   $12\ OH^-$ $\qquad$ $12\ OH^-$

   $\mathbf{6\ H_2O + 2\ BrO_3^- + 10\ Br^- \rightarrow 6\ Br_2 + 12\ OH^-}$

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

3. $8e^- + 9\ H^+ + CH_3COO^- \rightarrow 2\ CH_4 + 2\ H_2O$
   $2H_2O + CH_3COO^- \rightarrow 2\ CO_2 + 7H^+ + 8e^-$

   $2\ H^+ + 2\ CH_3COO^- \rightarrow 2\ CH_4 + 2\ CO_2$
   $2\ OH^-$ $\qquad$ $2\ OH^-$

   $\mathbf{2\ H_2O + 2\ CH_3COO^- \rightarrow 2\ CH_4 + 2\ CO_2 + 2\ OH^-}$

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

6.2 Activity Key

1. $\overset{+2}{C}N^- + 2\ \overset{+5}{Cl}O_3^- \rightarrow \overset{+4}{C}NO^- + \overset{0}{Cl_2}$ $\quad \Delta$ ON of C = +2, of Cl = (-5 x 2) = -10

$$\mathbf{2H^+ + 5\,CN^- + 2\,ClO_3^- \rightarrow 5\,CNO^- + Cl_2 + H_2O}$$

Species	Reactants	Products
Hydrogen	2	2
Nitrogen	5	5
Oxygen	6	6
Carbon	5	5
Chlorine	2	2
Charge	-5	-5

2. $\overset{0}{Fe} + \overset{+3}{As_2O_3} \rightarrow 2\,\overset{-3}{AsH_3} + Fe^{3+}$ Δ ON of Fe = +3, of As = (-6 x 2) = -12

$$12\,H^+ + 4\,Fe + As_2O_3 \rightarrow 2\,AsH_3 + 4\,Fe^{3+} + 3\,H_2O$$
$$12\,OH^- \qquad\qquad\qquad\qquad\qquad\qquad 12\,OH^-$$

$$\mathbf{9\,H_2O + 4\,Fe + As_2O_3 \rightarrow 2\,AsH_3 + 4\,Fe^{3+} + 12\,OH^-}$$

Species	Reactants	Products
Hydrogen	18	18
Arsenic	2	2
Oxygen	12	12
Iron	4	4
Charge	0	0

6.2 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. $4\ e^- + 6\ H^+ + CrO_4^{2-} \rightarrow Cr(OH)_2 + 2\ H_2O$
$(+\ 6\ OH^-)$ $(+6\ OH^-)$

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

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

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

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

5)

a. $3 \times (Sn^{2+} \rightarrow Sn^{4+} + 2\ e^-)$

$2 \times (3\ e^- + 4\ H^+ + MnO_4^- \rightarrow MnO_2 + 2\ H_2O)$

$8\ H^+ + 2\ MnO_4^- + 3\ Sn^{2+} \rightarrow 3\ Sn^{4+} + 2\ MnO_2 + 4\ H_2O$
$(+\ 8\ OH^-)$ $(+8\ OH^-)$

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

b. $4 \times (V^{2+} \rightarrow V^{3+} + e^-)$

$4\ e^- + 2\ H^+ + 2\ H_2SO_3 \rightarrow S_2O_3^{2-} + 3\ H_2O$

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

c. $10\ e^- + 12\ H^+ + 2\ IO_3^- \rightarrow I_2 + 6\ H_2O$

$5 \times (2\ I^- \rightarrow I_2 + 2\ e^-)$

$12\ H^+ + 2\ IO_3^- + 10\ I^- \rightarrow 6\ I_2 + 6\ H_2O$
$(+12\ OH^-)$ $(+12\ OH^-)$

$(6\ H_2O + 2\ IO_3^- + 10\ I^- \rightarrow 6\ I_2 + 12\ OH^-) \div 2$
$3\ H_2O + IO_3^- + 5\ I^- \rightarrow 3\ I_2 + 6\ OH^-$

d. $4 \times (6\ e^- + 6\ H^+ + ClO_3^- \rightarrow Cl^- + 3\ H_2O)$

$3 \times (2\ H_2O + N_2H_4 \rightarrow 2\ NO + 8\ H^+ + 8\ e^-)$

$\mathbf{3\ N_2H_4 + 4\ ClO_3^- \rightarrow 4\ Cl^- + 6\ NO + 6\ H_2O}$

e. $2 \times (3\ e^- + 4\ H^+ + NO_3^- \rightarrow NO + 2\ H_2O)$

$3 \times (Zn \rightarrow Zn^{2+} + 2\ e^-)$

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

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

f. $8 \times (H_2O + ClO_3^- \rightarrow ClO_4^- + 2\ H^+ + 2\ e^-)$

$16\ e^- + 18\ H^+ + 3\ ClO_3^- \rightarrow Cl^- + Cl_2 + 9\ H_2O$

$\mathbf{2\ H^+ + 11\ ClO_3^- \rightarrow 8\ ClO_4^- + Cl^- + Cl_2 + H_2O}$

g. $16 \times (3\ e^- + 4\ H^+ + MnO_4^- \rightarrow MnO_2 + 2\ H_2O)$

$3 \times (9\ H_2O + SnS_3O_3 \rightarrow 3\ SO_4^{2-} + Sn^{4+} + 18\ H^+ + 16\ e^-)$

$48\ e^- + 64\ H^+ + 16\ MnO_4^- \rightarrow 16\ MnO_2 + 32\ H_2O$

$27\ H_2O + 3\ SnS_3O_3 \rightarrow 9\ SO_4^{2-} + 3\ Sn^{4+} + 54\ H^+ + 48\ e^-$

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

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

h. $20\ e^- + 20\ H^+ + 2\ Mg_3(AsO_4)_2 + 6\ SiO_2 \rightarrow As_4 + 6\ MgSiO_3 + 10\ H_2O$

$10 \times (H_2O + C \rightarrow CO + 2\ H^+ + 2\ e^-)$

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

6) Balance

a. +4-----ΔON= (+4)------0
-1--ΔON= (-1)x2=(-2)--0

$SeO_3^{2-} + 2F^- \rightarrow Se + F_2$

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

b. +7 -----ΔON= (-3)------ +4

+3---ΔON= (+2)x2=(+4)----+5

$ReO_4^- + Sb_2O_3 \rightarrow ReO_2 + Sb_2O_5$

$4\ H^+ + 4\ ReO_4^- + 3\ Sb_2O_3 \rightarrow 4\ ReO_2 + 3\ Sb_2O_5 + 2\ H_2O$

$(+4\ OH^-)$ $(+4\ OH^-)$

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

c. 0 --------ΔON=(+4)------- +4

+5 -------ΔON=(-3)--- +2

$Pd + I^- + NO_3^- \rightarrow PdI_6^{2-} + NO$

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

d. +2 -----ΔON= (+2)---------- +4

+1---ΔON= (-2) ----- -1

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

$2\ H^+ + Pb(OH)_4^{2-} + BrO^- \rightarrow PbO_2 + Br^- + 3\ H_2O$

$(+2\ OH^-)$ $(+2\ OH^-)$

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

7) $2\ K^+ + Cr_2O_7^{2-} + CH_3CH_2OH + H^+ + Cl^- \rightarrow CH_3COOH + K^+ + Cl^- + Cr^{3+} + 3\ Cl^- + H_2O$

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

$3 \times (H_2O + CH_3CH_2OH \rightarrow CH_3COOH + 4\ H^+ + 4\ e^-)$

$16\ H^+ + 2\ Cr_2O_7^{2-} + 3\ CH_3CH_2OH \rightarrow 3\ CH_3COOH + 4\ Cr^{3+} + 11\ H_2O$

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

8) $Na^+ + Cl^- + 2\ H^+ + SO_4^{2-} + MnO_2 \rightarrow 2\ Na^+ + SO_4^{2-} + Mn^{2+} + 2\ Cl^- + H_2O + Cl_2$

$2e^- + 4\ H^+ + MnO_2 \rightarrow Mn^{2+} + 2\ H_2O$

$2\ Cl^- \rightarrow Cl_2 + 2\ e^-$

$4\ H^+ + MnO_2 + 2\ Cl^- \rightarrow Mn^{2+} + 2\ H_2O + Cl_2$

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

6.3 Using the Standard Reduction Potential (SRP) Table to Predict Redox Reactions

Warm Up

1. For example: both are lists of equilibria, both contain numerical values, both show one type of chemical converting to another (both show only one type of reaction), reactants get stronger as you go up on the left side of both, products get stronger as you go down on the right side of both.
2. reductions; oxidizing agents
3. alkali metals
4. fluorine; halogens

Practice Problems – Determining Whether a Spontaneous Redox Reaction Will Occur

1. yes; $2I^- + Br_2 \rightarrow 2Br^- + I_2$
2. no; both are oxidizing agents
3. yes; $2Ag^+ + Sn \rightarrow Sn^{2+} + 2Ag$
4. no; I_2 isn't a strong enough oxidizing agent to spontaneously react with Cl^-

Practice Problems – Determining the Outcome of a Single Replacement Reaction

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

Quick Check

1. Sr
2. Fe^{3+}
3. I^-
4. Sn^{2+}
5. B^-
6. +0.80 V
7. +2.87 V

Practice Problems – Determining the Predominant Redox Reaction

1. $Sn^{4+} + Ni \rightarrow Sn^{2+} + Ni^{2+}$ (* Question revised from complimentary copy)
2. $3Cu^{2+} + 2Al \rightarrow 2Al^{3+} + 3Cu$
3. $2Cu^+ \rightarrow Cu^{2+} + Cu$
4. $Cu + Br_2 \rightarrow 2Br^- + Cu^{2+}$

Practice Problems – Determining a Chemical's Concentration via Redox Titration

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

   b. $0.0150\ L \times 0.0200\ \frac{mol}{L} = 3.00 \times 10^{-4}\ mol\ Cr_2O_7^{2-}$

   $3.00 \times 10^{-4}\ mol\ Cr_2O_7^{2-} \times \frac{6\ mol\ Fe^{2+}}{1\ mol\ Cr_2O_7^{2-}} = 1.80 \times 10^{-3}\ mol\ Fe^{2+}$

   $[FeCl_2] = [Fe^{2+}] = \frac{1.80 \times 10^{-3}\ mol\ Fe^{2+}}{0.0200\ L} = 0.0900\ M$

2. $1.00\ g\ Na_2C_2O_4 \times \frac{1\ mol\ Na_2C_2O_4}{134.0\ g\ Na_2C_2O_4} = 7.46 \times 10^{-3}\ mol\ Na_2C_2O_4$

   $7.46 \times 10^{-3}\ mol\ C_2O_4^{2-} \times \frac{2\ mol\ MnO_4^-}{5\ mol\ C_2O_4^{2-}} = 2.99 \times 10^{-3}\ mol\ MnO_4^-$

   $[KMnO_4] = [MnO_4^-] = \frac{2.99 \times 10^{-3}\ mol\ MnO_4^-}{0.0163\ L} = 0.183\ M$

3. a. $0.02625\ L \times 0.500\ \frac{mol}{L} = 1.3125 \times 10^{-2}\ mol\ Cr_2O_7^{2-}$

   $1.3125 \times 10^{-2}\ mol\ Cr_2O_7^{2-} \times \frac{3\ mol\ CH_3CH_2OH}{2\ mol\ Cr_2O_7^{2-}} = 1.969 \times 10^{-2}\ mol\ CH_3CH_2OH$

   $[CH_3CH_2OH] = \frac{1.969 \times 10^{-2}\ mol\ CH_3CH_2OH}{0.0100\ L} = 1.97\ M$

   b. $1.97\ \frac{mol}{L} \times 46.0\ \frac{g}{mol} \times \frac{1\ mL}{0.789\ g} = \frac{115\ mL\ CH_3CH_2OH}{L\ beverage} = 11.5\%$

Activity: Making an SRP Table

1.

Oxidizing Agents		**Reducing Agents**
$Q + e^-$	⇄	Q^-
$D^{3+} + e^-$	⇄	D^{2+}
$P + e^-$	⇄	P^-
$M^{2+} + 2e^-$	⇄	M
$L^+ + e^-$	⇄	L
$C^{2+} + 2e^-$	⇄	C

2. Q^-
3. C^{2+}

Review Questions

1. a. no
   b. yes

2. Either Fe^{3+}, Hg_2^{2+}, Ag^+, or Hg^{2+}

3. a. yes; $3Mg + 2Al^{3+} \rightarrow 3Mg^{2+} + 2Al$
   b. no
   c. yes; $Hg^{2+} + 2Ag \rightarrow Hg + 2Ag^+$

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. $Cu^{+} + Cr^{2+} \rightarrow Cr^{3+} + Cu$ (* Question revised from complimentary copy)
    c. $Cu^{2+} + Sn^{2+} \rightarrow Sn^{4+} + Cu$ (* Question revised from complimentary copy)

17. $Zn + 2Fe^{3+} \rightarrow 2Fe^{2+} + Zn^{2+}$

18. a. C^{3+}
    b. D

    SRP Table

    $C^{3+} + e^{-} \rightleftharpoons C^{2+}$
    $A^{2+} + 2e^{-} \rightleftharpoons A$
    $B^{2+} + 2e^{-} \rightleftharpoons B$
    $D^{+} + e^{-} \rightleftharpoons D$

19. $0.0328\ L \times 0.200\ \frac{mol}{L} = 6.56 \times 10^{-3}\ mol\ Cr_2O_7^{2-}$

    $6.56 \times 10^{-3}\ mol\ Cr_2O_7^{2-} \times \frac{6\ mol\ I^-}{1\ mol\ Cr_2O_7^{2-}} = 3.936 \times 10^{-2}\ mol\ I^-$

    $[KI] = [I^-] = \frac{3.936 \times 10^{-2}\ mol\ I^-}{0.0150\ L} = 2.62\ M$

20. $0.105\ g\ H_2C_2O_4 \cdot 2H_2O \times \frac{1\ mol\ H_2C_2O_4 \cdot 2H_2O}{126.0\ g\ H_2C_2O_4 \cdot 2H_2O} = 8.33 \times 10^{-4}\ mol\ H_2C_2O_4 \cdot 2H_2O$

    $8.33 \times 10^{-4}\ mol\ H_2C_2O_4 \times \frac{2\ mol\ MnO_4^-}{5\ mol\ H_2C_2O_4} = 3.33 \times 10^{-4}\ mol\ MnO_4^-$

    $[KMnO_4] = [MnO_4^-] = \frac{3.33 \times 10^{-4}\ mol\ MnO_4^-}{0.0186\ L} = 0.0179\ M$

21. a. $2Cr_2O_7^{2-} + C_2H_5OH + 16H^+ \rightarrow 4Cr^{3+} + 11H_2O + 2CO_2$

b.

$$0.01015\ L \times 0.0150\ \frac{mol}{L} = 1.5225 \times 10^{-4}\ mol\ Cr_2O_7^{2-}$$

$$1.5225 \times 10^{-4}\ mol\ Cr_2O_7^{2-} \times \frac{1\ mol\ C_2H_5OH}{2\ mol\ Cr_2O_7^{2-}} = 7.6125 \times 10^{-5}\ mol\ C_2H_5OH$$

$$7.6125 \times 10^{-5}\ mol\ C_2H_5OH \times \frac{46.0\ g\ C_2H_5OH}{1\ mol\ C_2H_5OH} = 3.50 \times 10^{-3}\ g\ C_2H_5OH$$

$$\frac{3.50 \times 10^{-3}\ g\ C_2H_5OH}{5.00\ g\ C_2H_5OH} = 0.0700\%$$ Legally impaired now, not before Sep.2010.

6.4 The Electrochemical Cell

Warm Up

1. ion
2. electron
3. 6.25×10^{18}

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

Practice Problems – Determining Standard Cell Potentials (Voltages)

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

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,

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.

Review Questions

1.

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

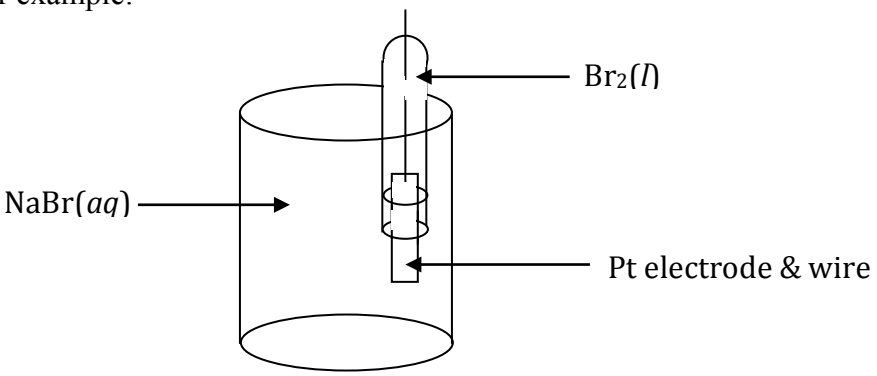
2. a. The diagram should show e^- flowing through the wire from the Fe/Fe^{2+} half-cell to the
   Ni/Ni^{2+} half-cell.
   b. Fe/Fe^{2+}
   c. $Fe \rightarrow Fe^{2+} + 2e^-$; $Ni^{2+} + 2e^- \rightarrow Ni$
   d. Fe labelled anode; Ni labelled cathode

e. $Fe \rightarrow Fe^{2+} + 2e^-$ $E° = +0.45$ V
$Ni^{2+} + 2e^- \rightarrow Ni$ $E° = -0.26$ V

$Fe + Ni^{2+} \rightarrow Fe^{2+} + Ni$ $E° = +0.19$ 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:

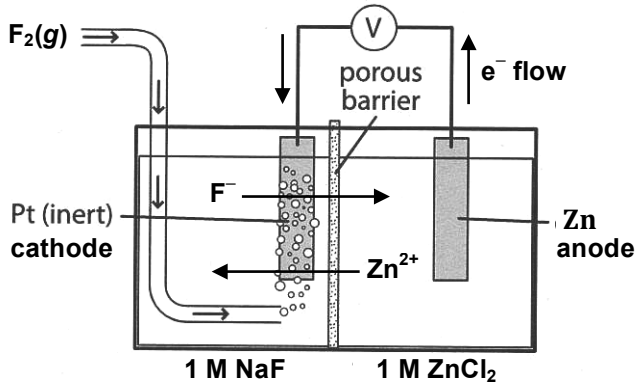


4. a. The diagram should show e^- flowing through the wire from the Cr/Cr^{3+} half-cell to the
H_2/H^+ half-cell
b. H_2/H^+
c. $Cr \rightarrow Cr^{3+} + 3e^-$; $2H^+ + 2e^- \rightarrow H_2$
d. Cr labelled anode; Pt labelled cathode
e. $2Cr \rightarrow 2Cr^{3+} + 6e^-$ $E° = 0.74$ V
$6H^+ + 6e^- \rightarrow 3H_2$ $E° = 0.00$ V

$2Cr + 6H^+ \rightarrow 2Cr^{3+} + 3H_2$ $E° = 0.74$ V

f. 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
g. The Cr electrode loses mass. The Pt electrode's mass is unchanged.

5.



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

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

6. it allows ion migration that completes the circuit

7. $2Fe^{2+} \rightarrow 2Fe^{3+} + 2e^-$ 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

8. 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 $Ag^+ + e^- \rightarrow Ag$)

9. $Cr^{2+} + 2e^- \rightarrow Cr$ 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

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

10. a. 0.83 V

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

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

b. –1.90 V

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

11.

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

Construct an <u>SRP Table</u> from the data provided:

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

12. 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

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

14. 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.

15. 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.

16. 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.

6.5 The Electrolytic Cell

Warm Up

Device	Energy Conversion	
	From	**To**
light bulb	electrical energy	light energy
photocell	light energy	electrical energy

speaker	electrical energy	**sound energy**
microphone	**sound energy**	electrical energy

heating element	electrical energy	**heat energy**
thermocouple	**heat energy**	electrical energy

motor	electrical energy	**mechanical energy**
generator	**mechanical energy**	electrical energy

electrolytic cell	electrical energy	chemical energy
electrochemical cell	**chemical energy**	electrical energy

Practice Problems – Predicting the Voltage Required to Operate an Electrolytic Cell

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

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

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

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

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

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

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

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

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

Quick Check

1. electro<u>chemical</u> cell
2. electro<u>lytic</u> cell
3. both
4. electro<u>chemical</u> cell
5. both
6. electro<u>lytic</u> cell
7. electro<u>chemical</u> cell

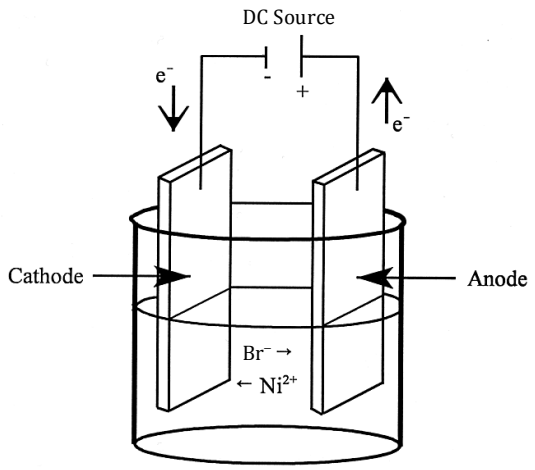
Practice Problems – Predicting the Half-Reactions That Will Occur in an Electrolytic Cell

1. a. Type 1 $2I^- \rightarrow I_2 + 2e^-$
$Mg^{2+} + 2e^- \rightarrow Mg$

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

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

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

Activity: Location, Location, Location

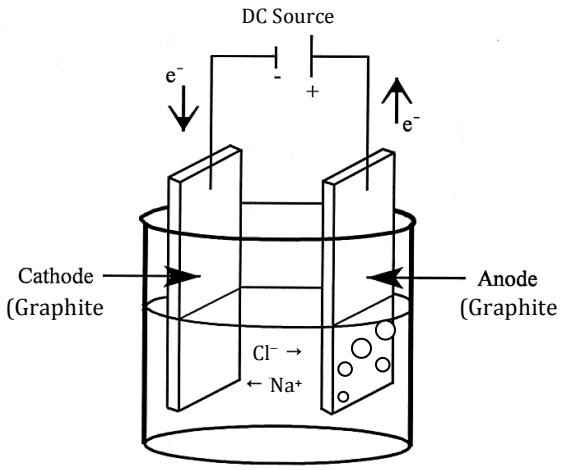
1. For example: proximity to source of ore
   proximity to shipping port
   proximity to market
   inexpensive electricity
   available work force
   inexpensive labour
   stable political climate
   democracy

3. ☒ Brazil
   ☑ deep water port
   ☒ western USA, Japan
   ☑ B.C. Hydro
   ☑ company town
   ☒ strong labour unions
   ☑ Canada – stable

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	**$½O_2 + 2H^+$**	**$H_2 + 2OH^-$**
e.	**3**	$CuCl_2$(*aq*)	Cu / Cu	**Cu^{2+}**	**Cu**

5. a. Anode (approximate) $H_2O \rightarrow ½\,O_2 + 2H^+ + 2e^-$ $E° = -1.37$ V

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.

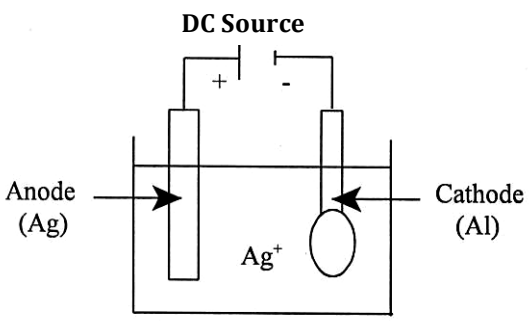
6. $2H_2O + 2e^- \rightarrow H_2 + 2\,OH^-$

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.



$$Ag \rightarrow Ag^+ + e^- \qquad E° = -0.80\ V$$
$$Ag^+ + e^- \rightarrow Ag \qquad E° = +0.80\ V$$
$$E° = 0.0$$

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

$$2Br^- \rightarrow Br_2 + 2e^- \qquad E° = -1.09\ V$$
$$Cu^{2+} + 2e^- \rightarrow Cu \qquad E° = +0.34\ V$$
$$2Br^- + Cu^{2+} \rightarrow Br_2 + Cu \qquad E° = -0.75\ V$$