

## Chapter 7

### Page 433, Practice Problems 7.1.1

- H = +1 O = -2
- Cs = +1 O = -1
- C = +4 O = -2
- Na = +1 Cr = +6 O = -2
- Ba = +2 H = -1
- N = -3 H = +1
- S = 0
- Al = +3 S = +6 O = -2

### Page 434, Quick Check

- 6 e<sup>-</sup>
- nothing (no e<sup>-</sup> transferred)

### Page 435, Quick Check

- You – oxidized (LEO)  
Partner - reduced (GER)
- Which partner acted as an “oxidizing agent”?  
my partner (they got reduced)  
A “reducing agent”? me (I got oxidized)
- 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

- $$0 \quad 0 \quad -4 +1$$
$$C(s) + 2 H_2(g) \rightarrow CH_4(g) \quad \text{O.A.} = C(s)$$
$$\text{R.A.} = H_2(g)$$
- $$0 \quad +3 -1 \quad 0 \quad +2 -1$$
$$3 Sr(s) + 2 FeBr_3(aq) \rightarrow 2 Fe(s) + 3 SrBr_2(aq)$$
$$\text{O.A.} = Fe^{3+} \quad \text{R.A.} = Sr$$
- $$+2 -2 \quad +5 -2 \quad +4 -2 \quad 0$$
$$5 CO(g) + Cl_2O_5(s) \rightarrow 5 CO_2(g) + Cl_2(s)$$
$$\text{O.A.} = Cl_2O_5 \quad \text{R.A.} = CO$$
- $$-3 +1 \quad 0 \quad 0$$
$$4 PH_3(g) \rightarrow P_4(g) + 6 H_2(g)$$
$$\text{O.A. and R.A.} = PH_3$$
- $$0 \quad +1 -2 \quad +2 -2 +1 \quad 0$$
$$Ba(s) + 2 H_2O(l) \rightarrow Ba(OH)_2(s) + H_2(g)$$
$$\text{O.A.} = H_2O \quad \text{R.A.} = Ba$$

### Page 438, 7.1 Review Questions

- 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.

- a) 
$$\begin{array}{cccc} \underline{Ca} & \underline{O} & \underline{C}_6 & \underline{Rb}_2 \\ +2 & +2 & 0 & -1 \end{array}$$
$$\underline{S}_2 & \underline{Be} & \underline{Br} & \underline{Cl} \\ +2 & -1 & +1 & 0$$
- a) A species that gets reduced/causes other species to be oxidized/gains e<sup>-</sup>  
b) A species that gets oxidized/causes other species to be reduced/loses e<sup>-</sup>  
c) Increased electronegativity makes a stronger O.A. and a weaker R.A.
- For each of the following reactions, indicate the species being oxidized and reduced and show the oxidation states above their symbols.
- $$+1 +5 -2 \quad +1 -1 \quad 0$$
$$2 KBrO_3(s) \rightarrow 2 KBr(s) + 3 O_2(g)$$
$$\text{Oxidized: } KBrO_3 \quad \text{Reduced: } KBrO_3$$
- $$0 \quad +1 +5 -2 \quad +2 +5 -2 \quad 0$$
$$Sr(s) + 2 CuNO_3(aq) \rightarrow Sr(NO_3)_2(aq) + 2 Cu(s)$$
$$\text{Ox'd: } Sr \quad \text{Red'd: } Cu^+$$
- $$0 \quad 0 \quad +2 -1$$
$$2 F_2(g) + O_2(g) \rightarrow 2 OF_2(g)$$
$$\text{Oxidized: } O_2 \quad \text{Reduced: } F_2$$
- $$-3 +1 +5 -2 \quad +1 -2 \quad +1 -2$$
$$NH_4NO_3(s) \rightarrow N_2O(g) + 2 H_2O(l)$$
$$\text{Ox'd: } NH_4NO_3 \quad \text{Red'd: } NH_4NO_3 \text{ (solid)}$$
- 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.
- $$0 \quad 0 \quad +2 -2$$
$$2 Sn(s) + O_2(g) \rightarrow 2 SnO(s)$$
$$\text{OA: } O_2 \quad \text{RA: } Sn \quad \text{No e}^-: 2$$
- $$0 \quad 0 \quad +5 -1$$
$$2 V(s) + 5 I_2(g) \rightarrow 2 VI_5(s)$$
$$\text{OA: } I_2 \quad \text{RA: } V \quad \text{No e}^-: 5$$
- $$0 \quad +1 -1 \quad +2 -1 \quad 0$$
$$Sr(s) + 2 HCl(aq) \rightarrow SrCl_2(aq) + H_2(g)$$
$$\text{OA: } H^+ \quad \text{RA: } Sr \quad \text{No e}^-: 2$$
- $$-8/3 +1 \quad 0 \quad +4 -2 \quad +1 -2$$
$$C_3H_8(g) + 5 O_2(g) \rightarrow 3 CO_2(g) + 4 H_2O(g)$$
$$\text{OA: } O_2 \quad \text{RA: } C_3H_8 \quad \text{No e}^-: 6\frac{2}{3}$$
- Write a balanced redox equation (in net ionic form) to show what has occurred in the

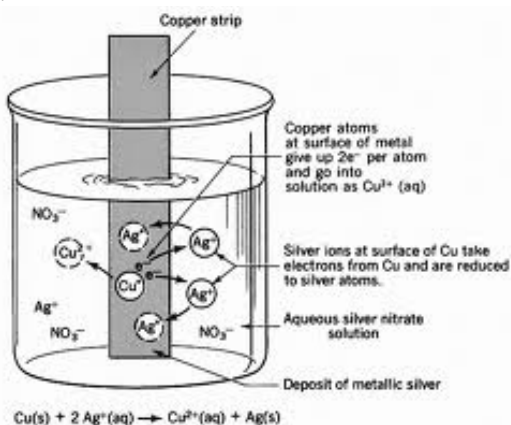
beaker over time:  $\text{Fe}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Fe}^{2+}(aq) + \text{Cu}(s)$

- 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)  $\text{P}^{3-}$  -3 b)  $(\text{NH}_4)_2\underline{\text{Zr}}(\text{SO}_4)_3 + 4$  c)  $\text{Na}_2\underline{\text{C}}_2\text{O}_4 + 3$   
 d)  $\underline{\text{N}}_2\text{H}_5\text{Cl} -2$  e)  $\underline{\text{Mn}}\text{O}_4^{2-} + 6$

8.



- a) see above  
 b) Oxidized? Cu metal  
 Reduced?  $\text{Ag}^+$  ion  
 c) Reducing agent? Cu metal  
 Oxidizing agent?  $\text{Ag}^+$  ion  
 d)  $2e^-$   
 a) Alkali metal family (IA)  
 b) Halogens (VIIA) or (17)

9. a)  $\text{ClO}_4^-$ ,  $\text{ClO}_3^-$ ,  $\text{ClO}_2^-$   
 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.  $2e^- + \text{H}_2\text{O} + \text{IO}_4^- \rightarrow \text{IO}_3^- + 2\text{OH}^-$   
 reduction  
 4.  $5 \text{H}_2\text{O} + \text{S}_2\text{O}_3^{2-} \rightarrow 2 \text{SO}_4^{2-} + 10 \text{H}^+ + 8e^-$   
 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.  $4 \text{Zn} + 8 \text{H}^+ + \text{H}_3\text{AsO}_4 \rightarrow \text{AsH}_3 + 4 \text{H}_2\text{O} + 4 \text{Zn}^{2+}$

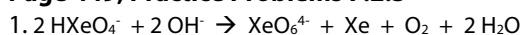
Species	Reactants	Products
Hydrogen	11	11
Zinc	4	4
Oxygen	4	4
Arsenic	1	1
Charge	+8	+8

2.

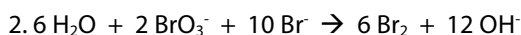


Species	Reactants	Products
Hydrogen	10	10
Nitrogen	4	4
Oxygen	13	13
Carbon	2	2
Charge	0	0

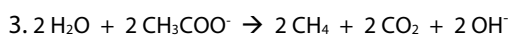
#### Page 449, Practice Problems 7.2.3



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



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



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 <sup>-</sup> to products
a.	3+	2-	5 e <sup>-</sup> to reactants
b.	1-	3-	2 e <sup>-</sup> to reactants
c.	2-	4+	6 e <sup>-</sup> to

			products
d.	1+	5+	4 e <sup>-</sup> 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)  $4 H_2O + 2 NO_3^- + 3 Zn \rightarrow 3 Zn^{2+} + 2 NO + 8 OH^-$
- 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)  $6 H^+ + SeO_3^{2-} + 4 F^- \rightarrow Se + 2 F_2 + 3 H_2O$
- 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)  $4 NaCl + 2 H_2SO_4 + MnO_2 \rightarrow 2 Na_2SO_4 + MnCl_2 + 2 H_2O + Cl_2$

#### Page 459, Practice Problems 7.3.1

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

#### Page 460, Practice Problems 7.3.2

- no;  $I_2$  isn't a strong enough oxidizing agent to spontaneously react with  $F^-$
- yes;  $3 Cu^{2+} + 2 Al \rightarrow 2 Al^{3+} + 3 Cu$
- yes;  $Cl_2 + 2 Br^- \rightarrow Br_2 + 2 Cl^-$  (\* Question revised from complimentary copy)
- no;  $Al^{3+}$  isn't a strong enough oxidizing agent to spontaneously react with Sn

#### Page 462, Quick Check

- Sr
- $Fe^{3+}$
- $I^-$
- $Sn^{2+}$
- $B^-$
- +0.80 V
- +2.87 V

#### Page 464, Practice Problems 7.3.3

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

#### Page 465, Practice Problems 7.3.4

- a)  $Cr_2O_7^{2-} + 6 Fe^{2+} + 14 H^+ \rightarrow 2 Cr^{3+} + 6 Fe^{3+} + 7 H_2O$   
b) 0.0900 M
- 0.183 M
- a) 1.97 M  
b) 11.5%

#### Page 467, 7.3 Review Questions

- a) no  
b) yes
- Either  $Fe^{3+}$ ,  $Hg_2^{2+}$ ,  $Ag^+$ , or  $Hg^{2+}$
- a) yes;  $3 Mg + 2 Al^{3+} \rightarrow 3 Mg^{2+} + 2 Al$   
b) no  
c) yes;  $Hg^{2+} + 2 Ag \rightarrow Hg + 2 Ag^+$

4.

Metals	Non-metals
bottom right of SRP Table	<b>top left of SRP Table</b>
<b>tend to give electrons</b>	tend to take electrons
give e <sup>-</sup> to chemicals above them on the left	take e <sup>-</sup> from chemicals below them on the right

5. a) yes;  $\text{Fe} + \text{Sn}^{2+} \rightarrow \text{Fe}^{2+} + \text{Sn}$   
 b) yes;  $\text{F}_2 + 2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{F}^-$   
 c) no
6. a)  $2\text{I}^- + 2\text{Fe}^{3+} \rightarrow \text{I}_2 + 2\text{Fe}^{2+}$   
 b)  $\text{Br}_2 + 2\text{Fe}^{2+} \rightarrow 2\text{Br}^- + 2\text{Fe}^{3+}$
7.  $3\text{Sn}^{2+} + 2\text{NO}_3^- + 8\text{H}^+ \rightarrow 2\text{NO} + 4\text{H}_2\text{O} + 3\text{Sn}^{4+}$
8. No. The  $\text{Fe}^{3+}$  would oxidize and dissolve the Al container.
9. a) reverse  
 b) forward
10.  $\text{Mg} + 2\text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2$
11. The  $\text{H}^+$  in acids is not a strong enough oxidizing agent to oxidize silver but the nitrate ion in nitric acid is.
12.  $2\text{Na} + \text{H}_2\text{O} \rightarrow 2\text{Na}^+ + \text{H}_2 + \text{OH}^-$   
 The  $\text{OH}^-$  turns the phenolphthalein pink. The  $\text{H}_2$  is ignited by the energy released by the exothermic reaction.
13. a)  $\text{Ag}^+$   
 b) Mg
14.  $\text{Ga}^{3+}$ ; Al is being oxidized by gallium ions therefore  $\text{Ga}^{3+}$  is a stronger oxidizing agent (higher on the left in the SRP table) than  $\text{Al}^{3+}$

#### 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  $[\text{Fe}^{2+}]$  decreases, Q decreases, log Q decreases, E increases.  
 $E = +0.23 \text{ V}$
2.  $E = 0.52 \text{ V}$
3.  $Q = 3.07 \times 10^{11} = 10^{11}$   
 $[\text{Cu}^{2+}] = \frac{1.0 \text{ M}}{10^{11}} = 10^{-12} \text{ M}$

and so  $\text{Ga}^{3+}$  has a greater reduction potential than  $\text{Al}^{3+}$ .

15. a) +1.09V  
 b) +0.76V
16. a)  $\text{Fe} + \text{Co}^{2+} \rightarrow \text{Fe}^{2+} + \text{Co}$   
 b)  $\text{Cu}^+ + \text{Cr}^{2+} \rightarrow \text{Cr}^{3+} + \text{Cu}$   
 c)  $\text{Cu}^{2+} + \text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + \text{Cu}$
17.  $\text{Zn} + 2\text{Fe}^{3+} \rightarrow 2\text{Fe}^{2+} + \text{Zn}^{2+}$
18. a)  $\text{C}^{3+}$   
 b) D
19. 2.62 M
20. 0.0179 M
21. a)  $2\text{Cr}_2\text{O}_7^{2-} + \text{C}_2\text{H}_5\text{OH} + 16\text{H}^+ \rightarrow 4\text{Cr}^{3+} + 11\text{H}_2\text{O} + 2\text{CO}_2$   
 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/ $\text{Cr}^{3+}$  half-cell to the Ag/ $\text{Ag}^+$  half-cell  
 4. Ag

#### Page 475, Practice Problems 7.4.1

1.  $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^- \quad E^\circ = 0.26 \text{ V}$   
 $\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^- \quad E^\circ = 1.09 \text{ V}$   


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 $\text{Br}_2 + \text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{Br}^- \quad E^\circ = 1.35 \text{ V}$
2.  $2\text{Al} \rightarrow 2\text{Al}^{3+} + 6\text{e}^- \quad E^\circ = 1.66 \text{ V}$   
 $3\text{Cu}^{2+} + 6\text{e}^- \rightarrow 3\text{Cu} \quad E^\circ = 0.34 \text{ V}$   


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 $2\text{Al} + 3\text{Cu}^{2+} \rightarrow 2\text{Al}^{3+} + 3\text{Cu} \quad E^\circ = 2.00 \text{ V}$
3.  $6\text{I}^- \rightarrow 3\text{I}_2 + 6\text{e}^- \quad E^\circ = -0.54 \text{ V}$   
 $2\text{Au}^{3+} + 6\text{e}^- \rightarrow 2\text{Au} \quad E^\circ = +1.50 \text{ V}$   


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 $6\text{I}^- + 2\text{Au}^{3+} \rightarrow 3\text{I}_2 + 2\text{Au} \quad E^\circ = +0.96 \text{ V}$

**Page 482, Quick Check**

1.

Type	Anode Material	Cathode Material	Electrolyte Medium	Use
Alkaline Cell	Zn powder packed around a brass pin	MnO <sub>2</sub> (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 <sub>2</sub>	H <sub>2</sub> SO <sub>4</sub> (aq)	automobiles
Fuel Cell	porous carbon	porous carbon	proton exchange membrane	e.g. electric vehicles, space travel,

**Page 484, Quick Check**

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

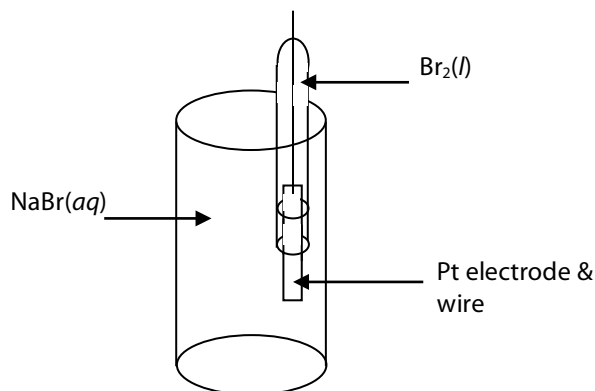
**Page 486, 7.4 Review Questions**

1.

Anode	Cathode
<b>oxidation occurs</b>	reduction occurs
mass decreases	<b>mass increases</b>
<b>attracts anions</b>	attracts cations
electrons flow away	<b>electrons flow towards</b>

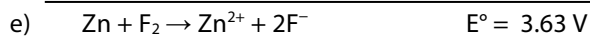
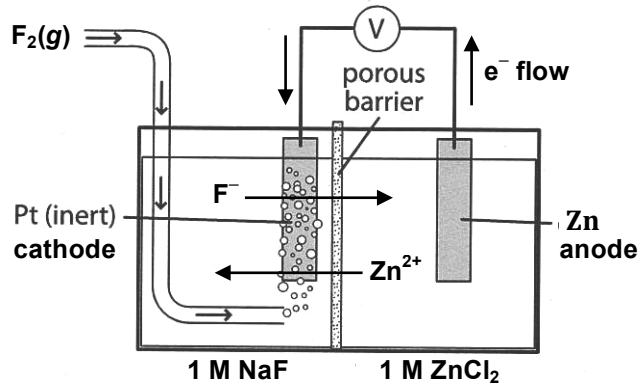
- The diagram should show e<sup>-</sup> flowing through the wire from the Fe/Fe<sup>2+</sup> half-cell to the Ni/Ni<sup>2+</sup> half-cell.
  - Fe/Fe<sup>2+</sup>
  - $\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^-$ ;  $\text{Ni}^{2+} + 2e^- \rightarrow \text{Ni}$
  - Fe labelled anode; Ni labelled cathode
  - E° = +0.19 V
  - NH<sub>4</sub><sup>+</sup> flow from the salt bridge into the Ni/Ni<sup>2+</sup> half-cell  
NO<sub>3</sub><sup>-</sup> flow through the salt bridge into the Fe/Fe<sup>2+</sup> half-cell
  - The Fe electrode loses mass. The Ni electrode gains mass.

3. For example:

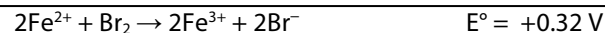


4. a) The diagram should show  $e^-$  flowing through the wire from the  $\text{Cr}/\text{Cr}^{3+}$  half-cell to the  $\text{H}_2/\text{H}^+$  half-cell  
 b)  $\text{H}_2/\text{H}^+$   
 c)  $\text{Cr} \rightarrow \text{Cr}^{3+} + 3e^-$ ;  $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$   
 5. a) Cr labelled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

6.

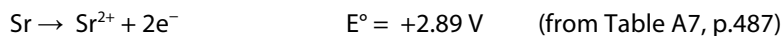
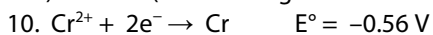


7. It allows ion migration that completes the circuit



9. a) 2.07 V (the  $E^\circ$  of  $\text{F}_2 + 2e^- \rightarrow 2\text{F}^-$  is 2.07 V greater than the  $E^\circ$  of  $\text{Ag}^+ + e^- \rightarrow \text{Ag}$ )

b) -3.17 V (the  $E^\circ$  of  $\text{Mg}^{2+} + 2e^- \rightarrow \text{Mg}$  is 3.17 V less than the  $E^\circ$  of  $\text{Ag}^+ + e^- \rightarrow \text{Ag}$ )



- $\text{Sr} + \text{Cr}^{2+} \rightarrow \text{Sr}^{2+} + \text{Cr} \quad E^\circ = +2.33 \text{ V} \quad (\text{provided by question})$   
 11a) 0.83 V  
 $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^- \quad E^\circ = -0.34 \text{ V} \quad (\text{from Table A7, p.487})$   
 $\text{Pd}^{2+} + 2\text{e}^- \rightarrow \text{Pd} \quad E^\circ = +0.83 \text{ V}$   


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 $\text{Cu} + \text{Pd}^{2+} \rightarrow \text{Cu}^{2+} + \text{Pd} \quad E^\circ = +0.49 \text{ V} \quad (\text{provided by question})$   
 b) -1.90 V  
 $2\text{Np} \rightarrow 2\text{Np}^{3+} + 6\text{e}^- \quad E^\circ = 1.90 \text{ V}$   
 $3\text{Pd}^{2+} + 6\text{e}^- \rightarrow 3\text{Pd} \quad E^\circ = 0.83 \text{ V} \quad (\text{answer to 10 a.})$   


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 $2\text{Np} + 3\text{Pd}^{2+} \rightarrow 2\text{Np}^{3+} + 3\text{Pd} \quad E^\circ = 2.73 \text{ V} \quad (\text{provided by question})$   
 1.90 V is the Standard Oxidation Potential (SOP) of Np therefore the Standard Reduction Potential (SRP) of  $\text{Np}^{3+}$  is -1.90 V.

12.

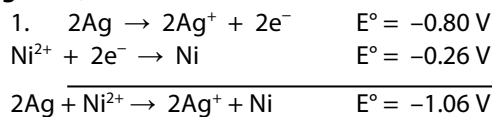
		Reduction Half-cell	
		$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt}$
Oxidation Half-cell	$\text{Pt} \rightarrow \text{Pt}^{2+} + 2\text{e}^-$	<b>- 1.60 V</b>	0 V
	$\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$	- 0.17 V	+ 1.43 V
	$\text{Ce} \rightarrow \text{Ce}^{3+} + 3\text{e}^-$	+ 1.93 V	<b>+ 3.53 V</b>

Construct an SRP Table from the data provided:

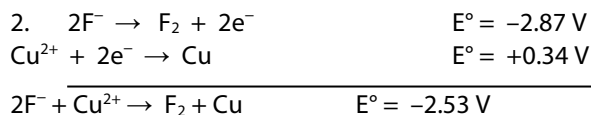
- $\text{Pt}^{2+} + 2\text{e}^- \rightleftharpoons \text{Pt} + 1.17 \text{ V} \quad (1.43 \text{ V above } \text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni})$   
 $\text{Ni}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ni} - 0.26 \text{ V} \quad (\text{from Table A7, p.487})$   
 $\text{Cd}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cd} - 0.43 \text{ V} \quad (0.17 \text{ V below } \text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni})$   
 $\text{Ce}^{3+} + 3\text{e}^- \rightleftharpoons \text{Ce} - 2.36 \text{ V} \quad (1.93 \text{ V below } \text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd})$
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  $\text{S}^{2-}$  ions precipitate out some  $\text{Fe}^{2+}$  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  $\text{CO}_2$  and  $\text{H}_2\text{O}$ . 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.  $E^\circ_{\text{cell}} = -0.86$ ;  $K = 10^{-150}$
20. a)  $E^\circ_{\text{cell}} = +2.20 \text{ V}$ ;  $K = 10^{223}$   
 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)  $E^\circ_{\text{cell}} = +0.76 \text{ V}$   
 b)  $7 \times 10^{-5} \text{ M}$   
 c)  $\text{pH} = 4.2$   
 22. a)  $E = +0.26 \text{ V}$   
 b)  $E^\circ_{\text{ox}} = +0.26 \text{ V}$   
 c)  $\text{Ni(s)}$

**Page 494, Practice Problems 7.5.1**



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



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

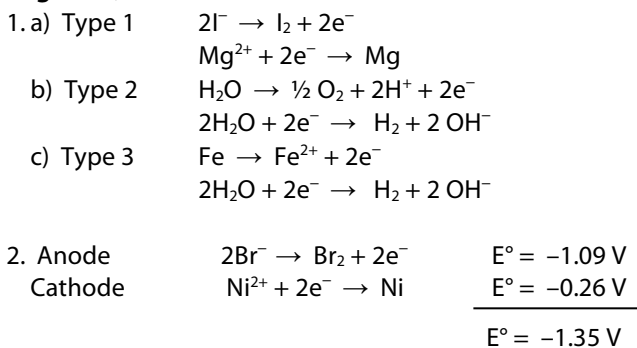


A voltage of at least 1.52 V would be required to operate this cell.

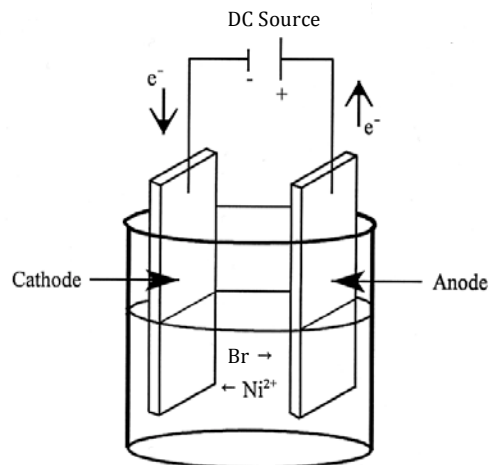
**Page 496, Quick Check**

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**



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



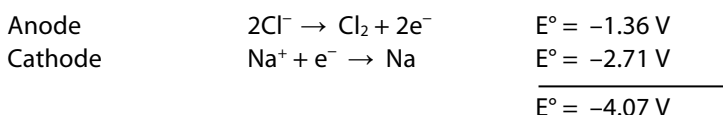
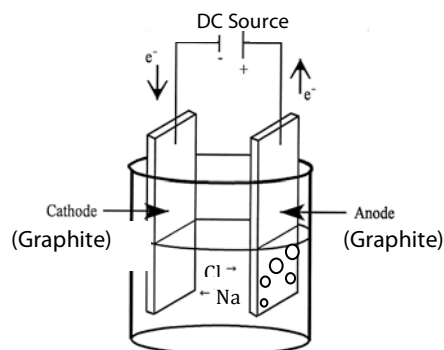


**Page 501, Quick Check**

1. the electrolytic recovery of a metal from a solution containing its ions.
2. aluminum oxide dissolves in molten cryolite ( $\text{Na}_3\text{AlF}_6(l)$ )
3. sodium hydroxide and chlorine
4. impressed current cathodic protection

**Page 503, 7.5 Review Questions**

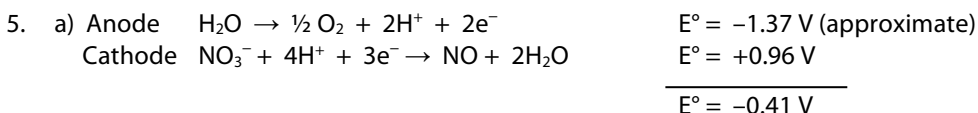
1. citric acid (weak electrolyte), sodium chloride, sodium citrate, and potassium dihydrogen phosphate
- 2.



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

3. Because the products ( $\text{Na}$  and  $\text{Cl}_2$ ) spontaneously react with each other
- 4.

	Cell Type (1,2,3)	Electrolyte	Anode/Cathode	Products	
				Anode	Cathode
a.	<b>1</b>	$\text{NaCl}(l)$	Pt / Pt	<b><math>\text{Cl}_2</math></b>	<b>Na</b>
b.	<b>2</b>	$\text{NaCl}(aq)$	Pt / C	<b><math>\text{Cl}_2</math></b>	<b><math>\text{H}_2 + 2\text{OH}^-</math></b>
c.	<b>2</b>	$\text{CuBr}_2(aq)$	C / C	<b><math>\text{Br}_2</math></b>	<b>Cu</b>
d.	<b>2</b>	$\text{AlF}_3(aq)$	C / C	<b><math>\frac{1}{2}\text{O}_2 + 2\text{H}^+</math></b>	<b><math>\text{H}_2 + 2\text{OH}^-</math></b>
e.	<b>3</b>	$\text{CuCl}_2(aq)$	Cu / Cu	<b><math>\text{Cu}^{2+}</math></b>	<b>Cu</b>

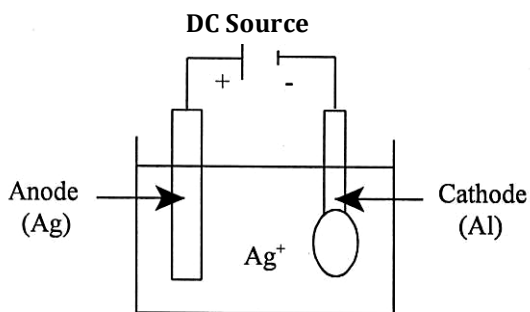


b) A voltage of at least 0.41 V would be required to operate this cell.

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



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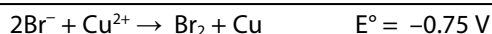

$$E^\circ = 0.0$$

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  $\text{H}_2\text{O}$ .

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

12. 0.75 V



### Page 506, Practice Problems

- 121 g Al (s)
- 1.5 A
- 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

- $\Delta G^\circ = -46 \text{ kJ/mol}_{rxn}$
- 0.40 V
- a) -2.19 V  
b)  $E^\circ_{ox} = -0.53 \text{ V}$   
 $\text{X}_2 = \text{I}_2 = \text{diatomic iodine}$   
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

- a) 0.44 g Cr (s)  
b) 820 s or 7.3 h  
c) 1.99 A
- 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<sub>2</sub> (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<sup>+</sup> 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)  $6 \text{ AlOF}_3^{2-} + 3 \text{ C(s)} \rightarrow 2 \text{ AlF}_6^{3-} + 3 \text{ CO}_2 + 4 \text{ Al}$   
b)  $\Delta G = 320. \text{ kJ/mol}_{\text{rxn}}$