$$n_{Ce^{4+}} = 15.1 \text{ mL} \times 0.125 \frac{\text{mol}}{\text{L}}$$

$$n_{Ce^{4+}} = 1.89 \text{ mmol}$$

$$n_{Fe^{2+}} = 1.89 \text{ mmol Ce}^{4+} \times \frac{1 \text{ mol Fe}^{2+}}{1 \text{ mol Ce}^{4+}}$$

$$n_{Fe^{2+}} = 1.89 \text{ mmol}$$

$$C_{Fe^{2+}} = \frac{1.89 \text{ mmol}}{25.0 \text{ mL}}$$

$$= 0.0755 \text{ mol/L}$$

$$C_{Fe^{2+}} = 75.5 \text{ mmol/L}$$

The concentration of iron(II) ions in the sample is 75.5 mmol/L.

# 9.3  PREDICTING REDOX REACTIONS

## PRACTICE

**(Page 676)**

### Understanding Concepts

1. Oxidation and reduction are processes that work together in the transfer of electrons that takes place in many chemical reactions. Oxidation is the process of losing electrons while reduction is the process of gaining electrons. Oxidizing agents and reducing agents are the substances directly involved in the transfer of electrons in a redox reaction. Oxidizing agents remove electrons from reducing agents, causing the reducing agent to be oxidized (loss of electrons) and the oxidizing agent to be reduced (gain of electrons).
2. If a substance is a very strong oxidizing agent it has a very strong attraction for electrons.
3. If a substance is a very strong reducing agent, its electrons are weakly attracted, and are easily removed.
4. According to Table 1, lead and zinc react spontaneously with a copper(II) ion solution.
5. According to Table 1, silver and copper did not appear to react with a copper(II) ion solution.
6. The metals, $Pb_{(s)}$ and $Zn_{(s)}$, which react spontaneously with $Cu^{2+}_{(aq)}$, both appear below the $Cu^{2+}_{(aq)}$ line in a table of reduction half-reactions.
7. Only zinc metal reacts spontaneously with lead(II) ions; silver, copper, and lead metals did not react with lead(II) ions. The metal, $Zn_{(s)}$, which reacts spontaneously with $Pb^{2+}_{(aq)}$, appears below the $Pb^{2+}_{(aq)}$ line in a table of reduction half-reactions.

### Applying Inquiry Skills

8. Metal ions react spontaneously with metals listed below them in a table of reduction half-reactions. According to Table 3, the predictions are correct and the hypothesis is verified.
9. The following table of reduction half-reaction equations is based on the evidence presented in Table 4.

**Relative Strengths of Oxidizing and Reducing Agents**

| decreasing | SOA | $Cl_{2(aq)} + 2 e^-$ | $\rightleftharpoons 2 Cl^-_{(aq)}$ | ↑ | decreasing |
|---|---|---|---|---|---|
| reactivity of | ↓ | $Br_{2(aq)} + 2 e^-$ | $\rightleftharpoons 2 Br^-_{(aq)}$ | ↑ | reactivity of |
| oxidizing agents | ↓ | $I_{2(aq)} + 2 e^-$ | $\rightleftharpoons 2 I^-_{(aq)}$ | SRA | reducing agents |

**PRACTICE**

## Understanding Concepts

10. $Co^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Co_{(s)}$

$Zn^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Zn_{(s)}$

$Mg^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Mg_{(s)}$

11. $Cu^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Cu_{(s)}$

$2\,H^+_{(aq)} + 2\,e^- \rightleftharpoons H_{2(g)}$

$Cd^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Cd_{(s)}$

$Be^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Be_{(s)}$

$Ca^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Ca_{(s)}$

12. The redox spontaneity rule is empirical because it is based directly on observations of redox reactions.

13. $Cl_{2(aq)} + 2\,e^- \rightleftharpoons 2\,Cl^-_{(aq)}$

$Br_{2(aq)} + 2\,e^- \rightleftharpoons 2\,Br^-_{(aq)}$

$Ag^+_{(aq)} + e^- \rightleftharpoons Ag_{(s)}$

$I_{2(aq)} + 2\,e^- \rightleftharpoons 2\,I^-_{(aq)}$

$Cu^{2+}_{(aq)} + 2\,e^- \rightleftharpoons Cu_{(s)}$

**PRACTICE**

## Understanding Concepts

14. $Ag^+_{(aq)}$, $Cu^{2+}_{(aq)}$, $Pb^{2+}_{(aq)}$, $Zn^{2+}_{(aq)}$. This order agrees with the evidence presented in Table 3 (text page 676).

15. (a) Metal ions, nonmetals, and a variety of acidic solutions usually behave as oxidizing agents.
    (b) Nonmetal ions, metals, and basic solutions of various entities usually behave as reducing agents.

16. According to atomic theory, nonmetal atoms have almost filled valence energy levels and tend to attract electrons to attain a noble gas-like electronic structure. Metal atoms, however, have few electrons in their valence energy levels, have weak attractions for the valence electrons, and tend to lose electrons to attain a noble gas-like electronic structure. This is consistent with the empirically determined table which shows that nonmetals tend to act as oxidizing agents (electron acceptors) and that metals tend to act as reducing agents (electron donors).

17. Since fluorine is the most reactive nonmetal, and nonmetals are generally oxidizing agents, fluorine is expected to be the strongest oxidizing agent. Fluorine is the most reactive nonmetal because it has the greatest attraction for electrons. This reason relates the observed reactivity to the theoretical definition of an oxidizing agent.

    *Some further "why" questions are as follows. Why does fluorine have the greatest attraction for electrons? Why does fluorine have the highest electronegativity? Why do atoms with almost-filled energy levels have higher electronegativities than those without? (It should soon become obvious that the theory presented, to this point, can provide only a limited explanation of the redox table.)*

18. Tin(II), copper(II), and chromium(II) ions can act as both oxidizing and reducing agents. These entities can gain or lose electrons. For example, tin(II) ions can either lose electrons to produce tin(IV) ions or gain electrons to produce tin atoms. The explanation for copper(II) and chromium(II) ions is similar, but there is no simple explanation for the behaviour of water molecules, which can also act as both oxidizing and reducing agents.

19. (a) spontaneous
    (b) nonspontaneous
    (c) nonspontaneous
    (d) spontaneous
    (e) spontaneous
    (f) spontaneous

## Applying Inquiry Skills

20. (I) Mix all combinations of oxidizing and reducing agents from a list of elements and their ions. The oxidizing agent that reacts spontaneously with the greatest number of reducing agents is the strongest. The oxidizing agent that reacts with the next greatest number of reducing agents is the next strongest, etc. From this information, the order of reactivity and the half-reaction equations can be determined.

   (II) Selected mixtures of oxidizing agents and reducing agents are studied and the spontaneity rule is used to order the oxidizing and reducing agents in each reactant mixture. The individual results are combined to form a final table of half-reactions. (If $n$ half-reaction equations are to be placed in a table, a minimum of $n-1$ reactions must be studied.)
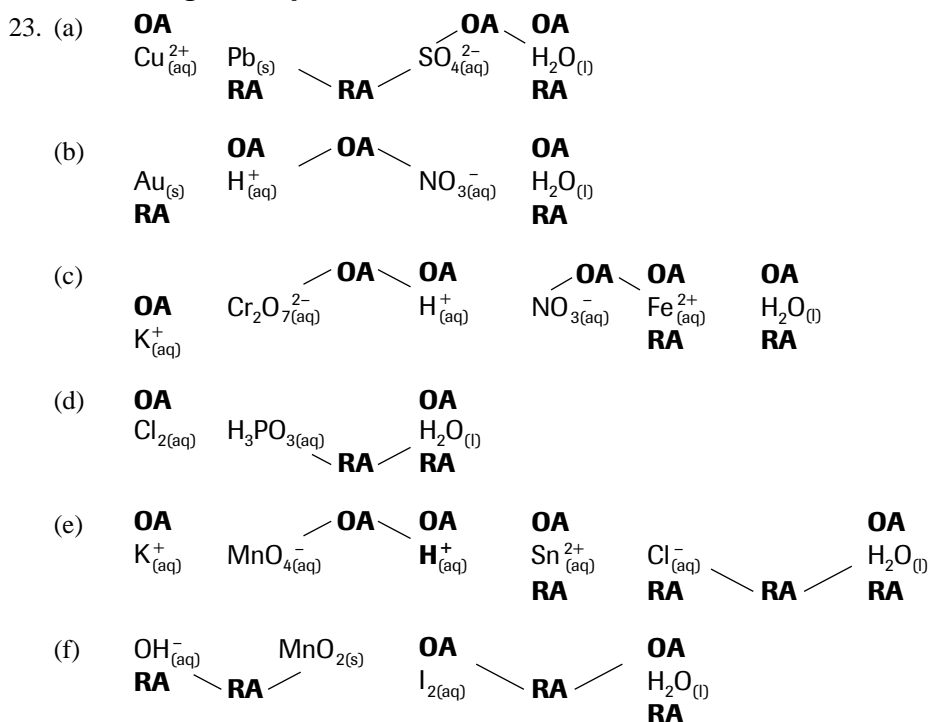
## Making Connections

21. Gold and silver (and platinum) are commonly found as metals, while sodium and potassium are never found in their elemental form. Gold and silver are listed near the top of the table of relative strengths of oxidizing and reducing agents and are very weak reducing agents (relatively unreactive). Sodium and potassium are listed near the bottom of the table of relative strengths of oxidizing and reducing agents and are very strong reducing agents (very reactive).

22. The empirical way of knowing has been most useful to this point in predicting the spontaneity of redox reactions because the construction of tables of relative strengths of oxidizing and reducing agents is based on observations of reactions. The theory presented to this point can barely provide an initial explanation after the fact, and has little predictive power.
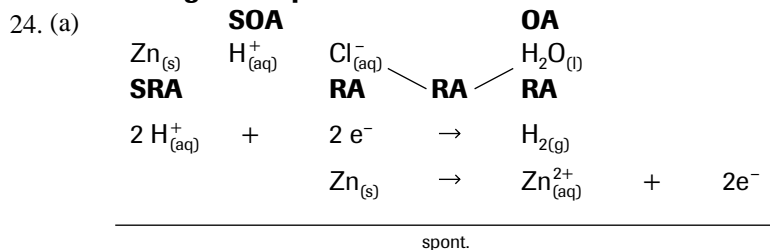
## PRACTICE

**(Page 680)**

### Understanding Concepts

23. (a) 

$$\text{OA: } Cu^{2+}_{(aq)} \quad Pb_{(s)} \quad SO_4^{2-}_{(aq)} \quad H_2O_{(l)}$$
$$\text{RA: } Pb_{(s)} \qquad \text{(OA, OA, RA)}$$

   (b)

$$Au_{(s)} \quad H^+_{(aq)} \quad NO_3^-_{(aq)} \quad H_2O_{(l)}$$

   (c)

$$K^+_{(aq)} \quad Cr_2O_7^{2-}_{(aq)} \quad H^+_{(aq)} \quad NO_3^-_{(aq)} \quad Fe^{2+}_{(aq)} \quad H_2O_{(l)}$$

   (d)

$$Cl_{2(aq)} \quad H_3PO_{3(aq)} \quad H_2O_{(l)}$$

   (e)

$$K^+_{(aq)} \quad MnO_4^-_{(aq)} \quad H^+_{(aq)} \quad Sn^{2+}_{(aq)} \quad Cl^-_{(aq)} \quad H_2O_{(l)}$$

   (f)

$$OH^-_{(aq)} \quad MnO_{2(s)} \quad I_{2(aq)} \quad H_2O_{(l)}$$

Each part (a)–(f) shows species labelled as OA (oxidizing agent) above and RA (reducing agent) below with connecting lines indicating spontaneous reactions.

## PRACTICE

**(Page 682)**

### Understanding Concepts

24. (a)

$$\begin{array}{cccc}
 & \textbf{SOA} & & \textbf{OA} \\
Zn_{(s)} & H^+_{(aq)} & Cl^-_{(aq)} & H_2O_{(l)} \\
\textbf{SRA} & \textbf{RA} & \textbf{RA} & \textbf{RA}
\end{array}$$

$$2\,H^+_{(aq)} \quad + \quad 2\,e^- \quad \rightarrow \quad H_{2(g)}$$
$$\qquad\qquad\qquad Zn_{(s)} \quad \rightarrow \quad Zn^{2+}_{(aq)} \quad + \quad 2e^-$$

$$\overline{\hspace{10cm}}$$
$$\text{spont.}$$
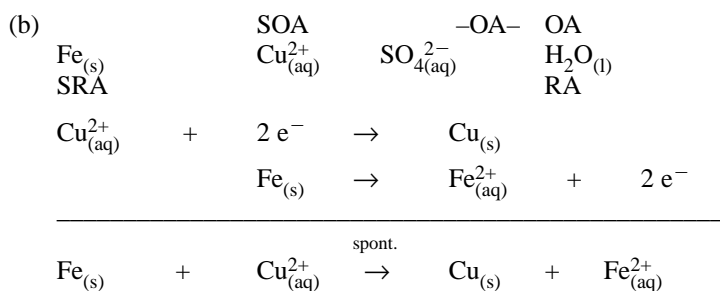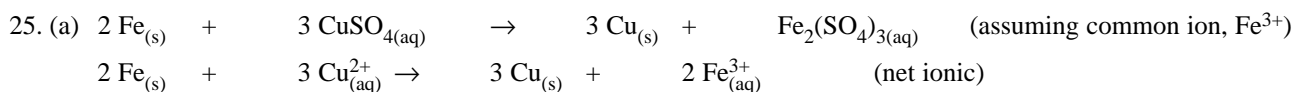
If a flame is inserted into a sample of the gas produced and a "pop" sound is heard, then hydrogen gas is likely present. (Other diagnostic tests include pre- and post-tests for pH, mass measurement of zinc, and a flame test for $Zn^{2+}_{(aq)}$.)

(b)

$$\begin{array}{cccc}
 & \textbf{SOA} & & \textbf{OA} \\
Au_{(s)} & H^+_{(aq)} & Cl^-_{(aq)} & H_2O_{(l)} \\
RA & RA & -RA- & SRA
\end{array}$$

$$2\,[2\,H^+_{(aq)} \quad + \quad 2\,e^- \quad \rightarrow \quad H_{2(g)}]$$
$$\qquad\qquad 2\,H_2O_{(l)} \quad \rightarrow \quad O_{2(g)} \;+\; 4\,H^+_{(aq)} \;+\; 4\,e^-$$

$$\overline{\hspace{10cm}}$$
$$\text{nonspont.}$$

$$2\,H_2O_{(l)} \quad \rightarrow \quad 2\,H_{2(g)} \quad + \quad O_{2(g)}$$

(c)

$$\begin{array}{cccc}
\text{OA} & -\text{SOA}- & & \text{OA} \\
H^+_{(aq)} & NO^-_{3(aq)} & Cu_{(s)} & H_2O_{(l)} \\
 & & SRA & RA
\end{array}$$

$$2\,[NO^-_{3(aq)} \quad + \quad 2\,H^+_{(aq)} \;+\; e^- \quad \rightarrow \quad NO_{2(g)} \quad + \quad H_2O_{(l)}\,]$$
$$\qquad\qquad\qquad\qquad\qquad Cu_{(s)} \quad \rightarrow \quad Cu^{2+}_{(aq)} \quad + \quad 2\,e^-$$

$$\overline{\hspace{10cm}}$$
$$\text{spont.}$$

$$2\,NO^-_{3(aq)} \;+\; 4\,H^+_{(aq)} \;+\; Cu_{(s)} \quad \rightarrow \quad 2\,NO_{2(g)} \;+\; 2\,H_2O_{(l)} \;+\; Cu^{2+}_{(aq)}$$
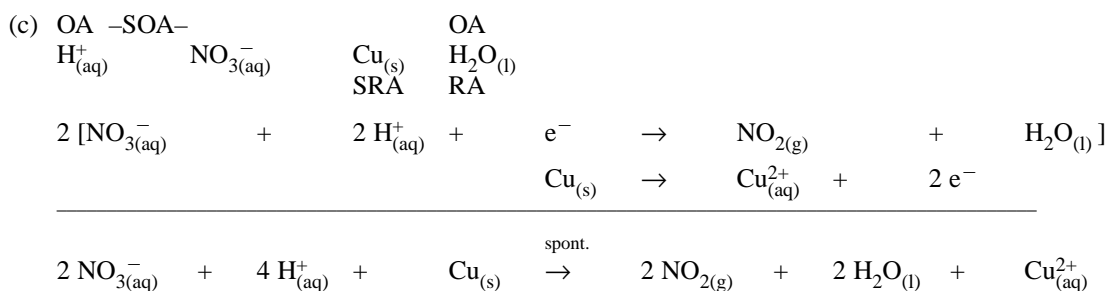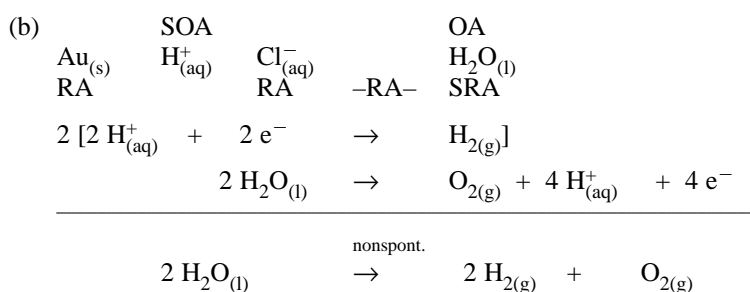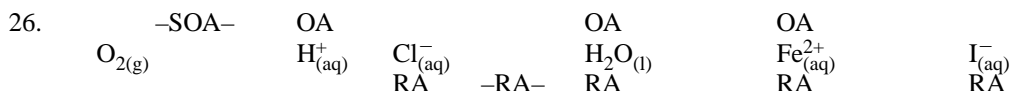
If the colour of the final solution near the copper surface is blue, then it is likely that copper(II) ions are produced. If a brown gas is produced during the course of the reaction, then nitrogen dioxide is likely produced. (Other diagnostic tests include pre- and post-tests for pH, mass measurement of copper, and a flame test for $Cu^{2+}_{(aq)}$.)

25. (a) $2\,Fe_{(s)} \quad + \quad 3\,CuSO_{4(aq)} \quad \rightarrow \quad 3\,Cu_{(s)} \;+\; Fe_2(SO_4)_{3(aq)}$ \quad (assuming common ion, $Fe^{3+}$)

$2\,Fe_{(s)} \quad + \quad 3\,Cu^{2+}_{(aq)} \rightarrow 3\,Cu_{(s)} \;+\; 2\,Fe^{3+}_{(aq)}$ \quad (net ionic)

(b)

$$\begin{array}{cccc}
 & \text{SOA} & -\text{OA}- & \text{OA} \\
Fe_{(s)} & Cu^{2+}_{(aq)} & SO^{2-}_{4(aq)} & H_2O_{(l)} \\
SRA & & & RA
\end{array}$$

$$Cu^{2+}_{(aq)} \quad + \quad 2\,e^- \quad \rightarrow \quad Cu_{(s)}$$
$$\qquad\qquad Fe_{(s)} \quad \rightarrow \quad Fe^{2+}_{(aq)} \quad + \quad 2\,e^-$$

$$\overline{\hspace{10cm}}$$
$$\text{spont.}$$

$$Fe_{(s)} \quad + \quad Cu^{2+}_{(aq)} \quad \rightarrow \quad Cu_{(s)} \;+\; Fe^{2+}_{(aq)}$$

(c) Both predictions cannot be correct; either iron(III) ions are formed or iron(II). The redox table and rules are more likely to be correct because these are based on extensive observations of relative strengths of oxidizing and
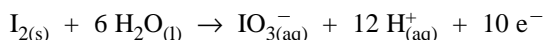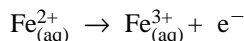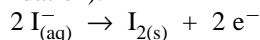
reducing agents supported by the idea of electron transfer. The single displacement rule is a generalization that has been useful in the past but does not have the same empirical basis or any theoretical justification.
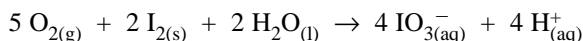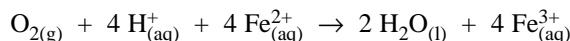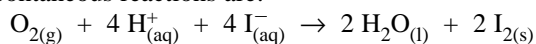
26.

$$\begin{array}{cccccc}
\text{–SOA–} & \text{OA} & & \text{OA} & \text{OA} & \\
O_{2(g)} & H^+_{(aq)} & Cl^-_{(aq)} & H_2O_{(l)} & Fe^{2+}_{(aq)} & I^-_{(aq)} \\
& & RA & \text{–RA–} \quad RA & RA & RA
\end{array}$$

Because an excess of acid is present, $O_{2(g)}$ and $H^+_{(aq)}$ will remain as the strongest oxidizing agent throughout all of the reactions.

$$O_{2(g)} \quad + \quad 4\,H^+_{(aq)} \quad + \quad 4\,e^- \quad \rightarrow \quad 2\,H_2O_{(l)}$$

Oxygen in an acidic solution will spontaneously oxidize (in order) iodide ions, iron(II) ions, and lastly iodine (product of the iodide oxidation).

$$2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$$

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-$$

$$I_{2(s)} + 6\,H_2O_{(l)} \rightarrow IO^-_{3(aq)} + 12\,H^+_{(aq)} + 10\,e^-$$

The spontaneous reactions are:

$$O_{2(g)} + 4\,H^+_{(aq)} + 4\,I^-_{(aq)} \rightarrow 2\,H_2O_{(l)} + 2\,I_{2(s)}$$

$$O_{2(g)} + 4\,H^+_{(aq)} + 4\,Fe^{2+}_{(aq)} \rightarrow 2\,H_2O_{(l)} + 4\,Fe^{3+}_{(aq)}$$

$$5\,O_{2(g)} + 2\,I_{2(s)} + 2\,H_2O_{(l)} \rightarrow 4\,IO^-_{3(aq)} + 4\,H^+_{(aq)}$$
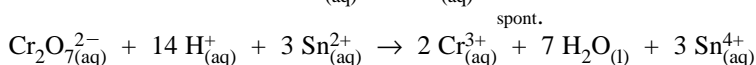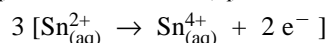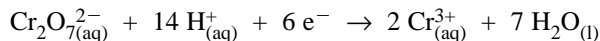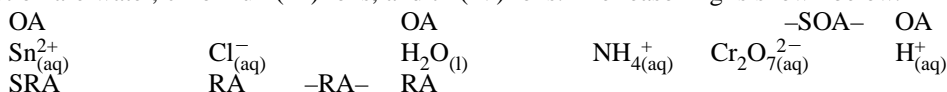
## Applying Inquiry Skills

27. Qualitatively, one could ensure complete reaction of the blue copper(II) solution and then observe the colour of the solution. If the iron(III) ion is produced, the solution should be light yellow, but if the iron(II) ion is produced, the solution should be light green.

Quantitatively, one could measure the mass of iron reacted and the mass of copper metal produced. Convert the masses to amounts and calculate the ratio of the amount of copper produced to the amount of iron reacted. If the ratio is 3:2, then iron(III) is produced, but if the ratio is 1:1, then iron(II) is produced.

28. (Sample response)

**Prediction**

According to redox concepts and the table of relative strengths of oxidizing and reducing agents, the products of the reaction are water, chromium(III) ions, and tin(IV) ions. The reasoning is shown below.

$$\begin{array}{cccccc}
\text{OA} & & \text{OA} & & \text{–SOA–} & \text{OA} \\
Sn^{2+}_{(aq)} & Cl^-_{(aq)} & H_2O_{(l)} & NH^+_{4(aq)} & Cr_2O^{2-}_{7(aq)} & H^+_{(aq)} \\
SRA & RA & \text{–RA–} \quad RA & & &
\end{array}$$

$$Cr_2O^{2-}_{7(aq)} + 14\,H^+_{(aq)} + 6\,e^- \rightarrow 2\,Cr^{3+}_{(aq)} + 7\,H_2O_{(l)}$$

$$3\,[Sn^{2+}_{(aq)} \rightarrow Sn^{4+}_{(aq)} + 2\,e^-\,]$$

$$\overset{\text{spont.}}{Cr_2O^{2-}_{7(aq)} + 14\,H^+_{(aq)} + 3\,Sn^{2+}_{(aq)} \rightarrow 2\,Cr^{3+}_{(aq)} + 7\,H_2O_{(l)} + 3\,Sn^{4+}_{(aq)}}$$
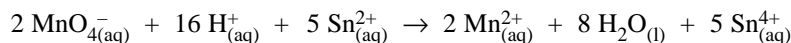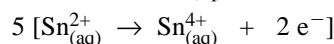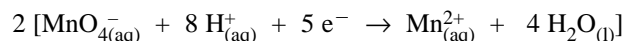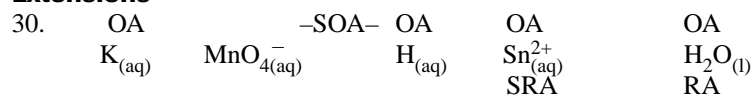
**Experimental Design**

An excess of tin(II) chloride is added to the acidic ammonium dichromate solution. If the colour of the mixture is observed before and after the reaction, and the reaction changes from orange to green, then chromium(III) ions are likely produced.

## Making Connections

29. Although aluminum is a reactive metal it does not corrode readily. When aluminum is exposed to air, a microscopically thin impermeable layer of aluminum oxide immediately forms on the surface, protecting it from further oxidation. The small pits that develop in aluminum cooking pots could be due to the reaction of aluminum with hydrogen ions, where the metal is exposed by scratching during stirring. The process is slow because the protective layer of aluminum oxide soon re-forms.

(**Note:** The presence of chloride ions (from salt) accelerates the pitting process, because chloride ions replace oxygens in the protective coating and reduce its protective integrity.)

## Extensions

30.

$$\text{OA} \qquad\qquad\qquad -\text{SOA}- \quad \text{OA} \qquad \text{OA} \qquad\qquad \text{OA}$$

$$K_{(aq)} \qquad MnO_{4(aq)}^{-} \qquad\quad H_{(aq)}^{+} \qquad Sn^{2+}_{(aq)} \qquad\qquad H_2O_{(l)}$$
$$\qquad\qquad\qquad\qquad\qquad\qquad\qquad \text{SRA} \qquad\qquad \text{RA}$$

$$2\,[MnO_{4(aq)}^{-} + 8\,H_{(aq)}^{+} + 5\,e^{-} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}]$$
$$5\,[Sn^{2+}_{(aq)} \rightarrow Sn^{4+}_{(aq)} + 2\,e^{-}]$$

$$2\,MnO_{4(aq)}^{-} + 16\,H_{(aq)}^{+} + 5\,Sn^{2+}_{(aq)} \rightarrow 2\,Mn^{2+}_{(aq)} + 8\,H_2O_{(l)} + 5\,Sn^{4+}_{(aq)}$$

12.4 mL                    10.00 mL

0.0832 mol/L         $C$

$$n_{MnO_4^{-}} = 12.4\ \text{mL} \times 0.0832\ \frac{\text{mol}}{\text{L}}$$

$$n_{MnO_4^{-}} = 1.03\ \text{mmol}$$

$$n_{Sn^{2+}} = 1.03\ \text{mmol}\,MnO_4^{-} \times \frac{5\ \text{mol Sn}^{2+}}{2\ \text{mol }MnO_4^{-}}$$

$$n_{Sn^{2+}} = 2.58\ \text{mmol}$$

$$C_{Sn^{2+}} = \frac{2.58\ \text{mmol}}{10.00\ \text{mL}}$$
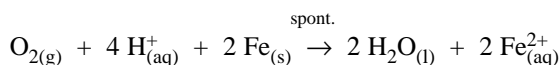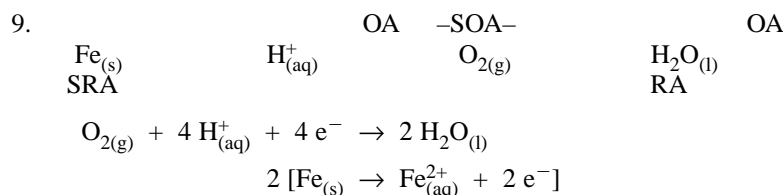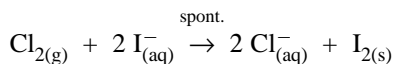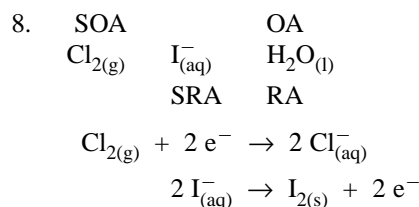
$$C_{Sn^{2+}} = 0.258\ \text{mol/L}$$

According to the evidence and stoichiometric analysis, the concentration of tin(II) ions is 0.258 mol/L.
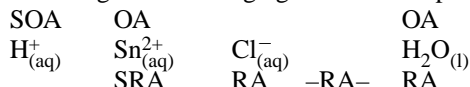
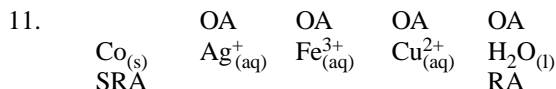## SECTION 9.3 QUESTIONS
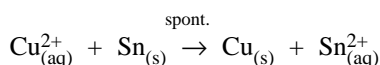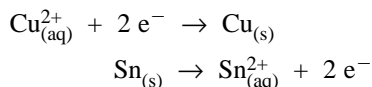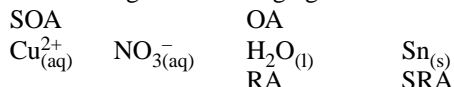
**(Page 683)**

### Understanding Concepts

1. The transfer of electrons between chemical entities is the key idea used to explain a redox reaction.
2. Oxidation is a process in which electrons are lost (LEO).
   Reduction is a process in which electrons are gained (GER).
3. Oxidation refers to the process of losing electrons, while oxidizing agent refers to the chemical entity that causes the oxidation of another entity by removing electrons from it.
4. Reduction refers to the process of gaining electrons, while reducing agent refers to the chemical entity that causes the reduction of another entity by donating electrons to it.
5. (a) reduction:        $Cu^{2+}_{(aq)} + 2\,e^{-} \rightarrow Cu_{(s)}$

         oxidation:        $Co_{(s)} \rightarrow Co^{2+}_{(aq)} + 2\,e^{-}$

   (b) reduction:        $Zn^{2+}_{(aq)} + 2\,e^{-} \rightarrow Zn_{(s)}$

         oxidation:        $Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + 2\,e^{-}$

   (c) reduction:        $Br_{2(l)} + 2\,e^{-} \rightarrow 2\,Br^{-}_{(aq)}$

         oxidation:        $2\,I^{-}_{(aq)} \rightarrow I_{2(s)} + 2\,e^{-}$

6. (a) spontaneous
   (b) nonspontaneous
   (c) spontaneous
7. $Pt^{4+}_{(aq)} + 4\,e^{-} \rightleftharpoons Pt_{(s)}$

   $2\,H^{+}_{(aq)} + 2\,e^{-} \rightleftharpoons H_{2(g)}$

   $Ni^{2+}_{(aq)} + 2\,e^{-} \rightleftharpoons Ni_{(s)}$

   $Ce^{3+}_{(aq)} + 3\,e^{-} \rightleftharpoons Ce_{(s)}$

   $Sr^{2+}_{(aq)} + 2\,e^{-} \rightleftharpoons Sr_{(s)}$

8.    SOA            OA

$Cl_{2(g)}$    $I^-_{(aq)}$    $H_2O_{(l)}$

            SRA    RA

$$Cl_{2(g)} + 2\,e^- \rightarrow 2\,Cl^-_{(aq)}$$

$$2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$$

----

$$Cl_{2(g)} + 2\,I^-_{(aq)} \overset{spont.}{\rightarrow} 2\,Cl^-_{(aq)} + I_{2(s)}$$

9.                                OA    –SOA–                    OA

$Fe_{(s)}$                $H^+_{(aq)}$           $O_{2(g)}$                $H_2O_{(l)}$

SRA                                               RA

$$O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^- \rightarrow 2\,H_2O_{(l)}$$

$$2\,[Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2\,e^-]$$

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$$O_{2(g)} + 4\,H^+_{(aq)} + 2\,Fe_{(s)} \overset{spont.}{\rightarrow} 2\,H_2O_{(l)} + 2\,Fe^{2+}_{(aq)}$$

10. (a) The solution will be stable, because no spontaneous reaction occurs according to the table of relative strengths of oxidizing and reducing agents and the spontaneity rule.

    SOA      OA                   OA

    $H^+_{(aq)}$     $Sn^{2+}_{(aq)}$    $Cl^-_{(aq)}$          $H_2O_{(l)}$
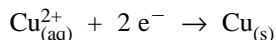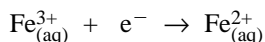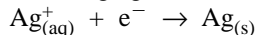
                     SRA      RA    –RA–    RA

(b) The solution will not be stable, because a spontaneous reaction occurs according to the table of relative strengths of oxidizing and reducing agents and the spontaneity rule.

    SOA                    OA

    $Cu^{2+}_{(aq)}$    $NO^-_{3(aq)}$    $H_2O_{(l)}$        $Sn_{(s)}$

                              RA          SRA

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

$$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2\,e^-$$

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$$Cu^{2+}_{(aq)} + Sn_{(s)} \overset{spont.}{\rightarrow} Cu_{(s)} + Sn^{2+}_{(aq)}$$

11.            OA      OA      OA      OA

    $Co_{(s)}$    $Ag^+_{(aq)}$    $Fe^{3+}_{(aq)}$    $Cu^{2+}_{(aq)}$    $H_2O_{(l)}$
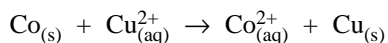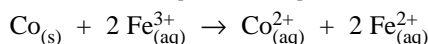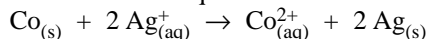
    SRA                                  RA

Because cobalt metal is in excess and it is the strongest reducing agent present, it will be the reducing agent in all reactions.
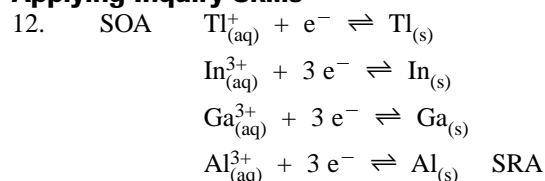
$$Co_{(s)} \rightarrow Co^{2+}_{(aq)} + 2\,e^-$$
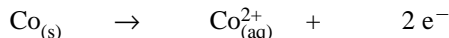
The oxidizing agents that will spontaneously react, in order of decreasing strength, are: $Ag^+_{(aq)}$, $Fe^{3+}_{(aq)}$, $Cu^{2+}_{(aq)}$.

$$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$$

$$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$$

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

The balanced redox equations are as follows:

$$Co_{(s)} + 2\,Ag^+_{(aq)} \rightarrow Co^{2+}_{(aq)} + 2\,Ag_{(s)}$$

$$Co_{(s)} + 2\,Fe^{3+}_{(aq)} \rightarrow Co^{2+}_{(aq)} + 2\,Fe^{2+}_{(aq)}$$

$$Co_{(s)} + Cu^{2+}_{(aq)} \rightarrow Co^{2+}_{(aq)} + Cu_{(s)}$$

## Applying Inquiry Skills

12.    SOA    $Tl^+_{(aq)} + e^- \rightleftharpoons Tl_{(s)}$

$In^{3+}_{(aq)} + 3\,e^- \rightleftharpoons In_{(s)}$

$Ga^{3+}_{(aq)} + 3\,e^- \rightleftharpoons Ga_{(s)}$

$Al^{3+}_{(aq)} + 3\,e^- \rightleftharpoons Al_{(s)}$    SRA

## Making Connections

13. Ursula Franklin specializes in the structure of metals and alloys and has pioneered the study of metallurgy in ancient cultures. In her view, science and technology have a side-by-side realtionship, stimulating and utilizing each other. She has voiced serious concerns about Canada's lack of support for research in science and technology. She is a tireless worker for Science for Peace, and has worked to improve opportunities for women in science and to make scientific information understandable and accessible to the public.

# 9.4  TECHNOLOGY OF CELLS AND BATTERIES

## PRACTICE

**(Page 687)**

### Understanding Concepts

1.  A simple electric cell contains two solid conductors (electrodes) and an aqueous conductor (electrolyte).
2.  An electrode is a solid conductor and an electrolyte is an aqueous conductor. An anode is the negative electrode in an electric cell. A cathode is the positive electrode in an electric cell.
3.  Four cells are needed. (6 V $\times$ 1.5 V/cell $=$ 4 cells)
4.  Electric current is a measure of the rate of flow of charge past a point in an electric circuit. Voltage is a measure of the energy difference per unit charge.

### Making Connections

5.  From a scientific perspective, a diagram shows where the electrons are believed to enter and exit the device.
    From a technological perspective, a diagram shows how the batteries should be placed in the device to make it work and to avoid damaging certain components.

## PRACTICE

**(Page 690)**

### Understanding Concepts

6. Scientific knowledge may be used to describe, explain, or predict the parts or operation of a device. Technological problem solving is used to develop a device that works based on established criteria.
7. The steps involved in technological problem solving are:
   * Develop a general design, for example, select variables to manipulate and control.
   * Follow several prediction–procedure–evidence–analysis cycles, manipulating and systematically studying one variable at a time.
   * Complete an evaluation on criteria such as efficiency, reliability, cost, and simplicity.
8. A technical and safety problem is the possible leakage of the liquid electrolyte. The electrolyte should be changed to a moist paste to reduce this problem.
9. Primary cells, such as the alkaline dry cell or mercury cell, cannot in practice be recharged.
   Secondary cells, such as the Ni–Cad cell, NiMH cell, or lead–acid cell, can be recharged repeatedly.
10. Advantages of the zinc chloride dry cell include simplicity, reliability, and relatively low cost. The main disadvantage is that the cell is not rechargeable and must be discarded when depleted. Other possible disadvantages for certain uses include short shelf life, relatively low current produced, and constancy of voltage.

### Making Connections

11. The AA, C, and D cells differ in size, current produced, and longevity of the cell, but the cell potential remains the same in each cell size. The smaller the cell, the smaller the electrodes and the smaller the current produced.
12. Both designs use containers that are consumed as they are used.