Chemistry 12

Solutions Manual Part A

Unit 5 Electrochemistry

Answers to Unit 5 Preparation Questions
(Student textbook pages 578-81)

1. a. corrosive material; precautions include wearing chemical safety goggles and a lab coat or apron in the laboratory, as well as other protective equipment, such as gloves, as directed. If possible, wear eyeglasses instead of contact lenses. (Eyeglasses are not a substitute for proper eye protection.)

b. flammable and combustible material; precautions include knowing the location and proper use of the nearest fire extinguisher, fire blanket, fire alarm, and drench hose/shower. Another precaution is to know what type of fire-fighting equipment should be used on particular types of fires. Additionally, tie back long hair and any loose clothing before starting an investigation. Lighters and matches must not be brought into the laboratory.

c. caustic chemicals that might irritate the skin or micro-organisms that might transmit infection; precautions include wearing a lab coat or apron and gloves to protect the skin. Additionally, do not wear open-toed shoes or sandals in the laboratory.

2. Student maps should show the location of at least three safety stations. Explanations of how to use equipment correctly should also discuss the circumstances in which the equipment is used.

3. a. Eyeglasses are not a substitute for proper eye protection, as chemicals may splash around the rims of the glasses into the eye.

b. If you get any material in your eyes, do not rub them. Wash your eyes immediately and continuously for 15 minutes at the nearest eyewash station, and make sure that your teacher is informed. A doctor should examine any eye injury. It is recommended not to wear contact lenses when working with chemicals. However, if you are wearing them, take them out prior to using the eyewash if possible. Failing to do so may result in the contact lenses doing further damage to your eyes or having material become trapped behind the contact lenses. Flush your eyes with water for 15 minutes, as above.

4. Answers may contain any three of the procedures below, as well as any other reasonable responses.

- Know the location and proper use of the nearest fire extinguisher, fire blanket, fire alarm, and drench hose/shower.
- Understand what type of fire extinguisher you have in the laboratory, and what type of fires it can be used on.
- Tie back long hair and any loose clothing before starting an investigation.
- Ties, scarves, long necklaces, and dangling earrings should be removed before starting an investigation.
- Organize materials and equipment neatly and logically. Place materials where you will not have to reach behind or over a Bunsen burner to get them.
- If your clothing catches fire, STOP, DROP, and ROLL. Other students may use the fire blanket to smother the flames. Do not wrap the fire blanket around yourself while in a standing position. This could result in a "chimney effect," bringing fire directly into your face.
- Do not look directly into the barrel of a Bunsen burner.
- If you see any of your classmates jeopardizing their safety or the safety of others while using a Bunsen burner, let your teacher know.
- When heating any item, wear safety eyewear, heat-resistant safety gloves, and any other safety equipment that your teacher or the Safety Precautions suggest.
- Always use heat-proof, intact containers. Check that there are no cracks in beakers or flasks.
- Never point the open end of a container, such as a test tube, that is being heated in a direction that could cause injury. The mouth of the container or test tube should point away from you and from others.
- Do not allow a container to boil dry unless specifically instructed to do so.
- Handle hot objects carefully. Be especially careful with a hot plate that may look as though it has cooled down, or glassware that has recently been heated.

- Before using a Bunsen burner, make sure that you understand how to light and operate it safely. Always pick it up by the base. Never leave a Bunsen burner unattended.
- Before lighting a Bunsen burner, make sure that there are no flammable solvents nearby.
- Always have the Bunsen burner secured to a utility stand.
- If you do receive a burn, run cold water over the burned area immediately. Make sure that your teacher is notified.

5. b

6. a. $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

b. $6Na(s) + Fe_2O_3(s) \rightarrow 3Na_2O(s) + 2Fe(s)$

c. $(NH_4)_2Cr_2O_7(s) \rightarrow N_2(g) + Cr_2O_3(s) + 4H_2O(g)$

7. a. reactants

b. products

c. equal

d. coefficients

8. a. $C_3H_8(g) + 5O_2(g) \rightarrow 4H_2O(g) + 3CO_2(g)$

b. $2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH(aq) + H_2(g)$

c. $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(\ell)$

9. The student's answer is incorrect. The correct answer is:
$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

10. a. NO_3^-

b. Consider a polyatomic ion as a single unit if it appears in the reactants and the products.

c. $2AgNO_3(aq) + CaCl_2(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$

11. a. decomposition reaction; a reaction in which a compound breaks down into two or more elements or simpler compounds

b. synthesis reaction; a reaction in which two or more substances react to produce a single compound

c. single displacement reaction; a reaction in which one element takes the place of another element in a compound

d. double displacement reaction; a reaction in which the positive ions of two ionic compounds exchange places, resulting in the formation of two new ionic compounds

12. decomposition reaction

13. single displacement reaction

14. a. The blue flame indicates that complete combustion is occurring. The yellow flame indicates that incomplete combustion is occurring. The reactions differ because the yellow flame is not getting enough oxygen for complete combustion to occur. The yellow colour is caused by the glow of carbon particles that is one product of incomplete combustion. The correct amount of oxygen results in complete combustion and too little oxygen results in incomplete combustion.

b. complete combustion:
$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
incomplete combustion:
$CH_4(g) + O_2(g) \rightarrow CO(g) + H_2O(g) + C(s)$

15. a. Students should predict that a single displacement reaction will occur. The products will be hydrogen gas and magnesium chloride.
$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

b. Students should predict that a synthesis reaction will occur. The product will be copper(II) chloride.
$Cu(s) + Cl_2(g) \rightarrow CuCl_2(s)$

c. Students should predict that a double displacement reaction will occur. The products will be lead chloride and lithium nitrate.
$2LiCl(aq) + Pb(NO_3)_2(aq) \rightarrow 2LiNO_3(aq) + PbCl_2(s)$

16. Copper pipes are resistant to corrosion, while steel pipes are not. The activity series of metals shows that copper is much less reactive than iron, which is a chief component of steel pipes.

17. a. The acid in acid precipitation is displaced by metals such as magnesium from the soil in a single displacement reaction. When an acid is a reactant in a single displacement reaction, you can think of the acid in terms of the ions it forms when it is dissolved in water: hydrogen ions and anions. The metal displaces the hydrogen and forms an ionic compound with the anion. The location of hydrogen in the activity series of metals shows the relative reactivity of hydrogen in acids. Every metal above hydrogen in the activity series can displace hydrogen from an acid, but the metals below hydrogen cannot. Magnesium is above hydrogen, so it can displace hydrogen to form an ionic compound with the anion. As an aqueous solution, this compound is subsequently leached from the soil.

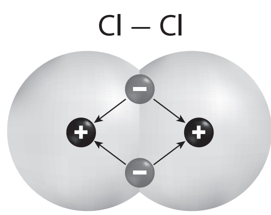
b. $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

18. Silver metal will not react with copper(II) nitrate because silver is below copper in the activity series of metals. As a result, it is less reactive than copper and will not displace it in a solution of copper(II) nitrate.

19. The symbol $\delta+$ indicates that the carbon atom is partially positively charged. The symbol $\delta-$ indicates that the chorine atom is partially negatively charged.

20. Captions should explain that if the electronegativity of one of two atoms that are bonded together is greater than the electronegativity of the other atom, the electrons will be attracted more strongly to nucleus of the more electronegative atom and spend more time there.

21. There are two chlorine atoms in a molecule of chlorine gas. Because the nuclei of both atoms in the molecule attract the shared electrons, the overall electronegativty for the molecule is zero.



chlorine molecule

22. The statement is incorrect. The noble gases do not share electrons with other atoms, so they do not have values for electronegativity.

23. a. 3.4

b. 1.0

c. 2.6

24. The electronegativity of the elements increases going up a group and going from left to right as the effective nuclear charge increases across a period in the periodic table. Thus, oxygen has a greater electronegativity than the other two atoms. It is above sulfur in Group 16 and to the right of lithium in Period 2.

25. a. *Same*: Both terms describe materials that allow different degrees of electron movement. *Different*: Conductors include metals and other materials in which electrons can easily move between atoms. Insulators include non-metal materials in which electrons cannot easily move from one atom to another.

b. *Same*: Both are voltaic cells that generate electric current by chemical reactions involving two different metals or metal compounds separated by an electrolyte. *Different*: A dry cell contains an electrolyte that is a paste. A wet cell contains an electrolyte that is a liquid.

c. *Same*: Both are voltaic cells that generate electric current by chemical reactions involving two different metals or metal compounds separated by an electrolyte which in each case may be aqueous (Daniell cell, lead storage battery) or a paste (dry cell, lithium battery). *Different*: A primary cell is a single cell which is not rechargeable; a secondary cell consists of two or more cells connected in series; the reaction in a secondary cell can be reversed and is therefore rechargeable.

d. *Same*: Both terms deal with electricity.

Different: An electric charge is the positive or negative charge on particles, such as electrons or protons. An electric current refers to the rate of movement of an electric charge. In an electric circuit, the current is due to the flow of electrons.

26. The electrodes play a role in generating electric current. The aluminum electrode gives up electrons that pass through the connecting wire to the copper electrode. The aluminum ion that is formed in this process enters the electrolyte solution. At the inert Cu electrode, $H^+(aq)$ accept electrons and $H_2(g)$ forms.

27. When atoms in the aluminum electrode give up electrons that move along the connecting wire to the copper electrode, ions from the aluminum electrode enter the electrolyte. Hydrogen ions in the electrolyte accept the electrons at the copper electrode. Thus, the electrolyte completes the circuit and enables an electric current to flow through the cell.

28. Different metals have different abilities to hold electrons. The electrodes in this cell must be made of different metals or metal compounds so that one can act as the electron donor and the other as the electron acceptor.

29. a. The aluminum electrode slowly disintegrates over time, as it loses aluminum atoms when it passes electrons to the connecting wire and aluminum ions enter the electrolyte solution.

b. It limits the life of the cell.

30. c

31. a. 1.5 V refers to how much work the cell can do. Each cell is able to do 1.5 J of work.

b. The cells are different sizes because they contain different quantities of materials, which affect how long they run. The D cell contains the most materials, so it will run the longest.

32. a. The light bulb requires a higher voltage than can be supplied by this cell.
   b. By adding more cells in series, the voltage can be increased so that it is large enough to light the bulb.

Chapter 9 Oxidation-Reduction Reactions

Learning Check Questions

(Student textbook page 586)

1. By definition, oxidation involves the loss of electrons, so a compound or atom can become oxidized if it loses electrons, without the need to combine with oxygen.
2. The original definition of reduction is the process of extracting metals from ores. The modern definition of reduction is the process of ions, atoms, or molecules gaining electrons. They are the same in that both definitions can refer to reactions involving metals.

   They differ in that the modern definition states that reduction involves a gain of electrons, and that materials other than metals can be reduced, whereas the original definition involves a loss of mass and refers only to metals.
3. Originally, oxidation was defined as any chemical reaction in which atoms or compounds react with molecular oxygen. The modern definition is the process in which an atom, ion, or molecule loses electrons.
4. Oxidation is the loss of electrons. These electrons must be accepted by another reagent. The ion/atom/molecule that gains the electrons undergoes reduction.
5. Oxidizing agent: $Fe^{2+}(aq)$; reducing agent: $Al(s)$
6. Using this reaction as a model for double displacement reactions, these reactions do not experience an exchange of electrons, and so double displacement reactions tend not to be redox reactions.

(Student textbook page 588)

7. a and c
8. Sample answer: zinc, barium, and magnesium.
9. You would use silver nitrate because silver ions are stronger oxidizing agents than cobalt ions and cobalt metal is a stronger reducing agent than silver metal. Under these conditions, the reaction will occur spontaneously.
10. $Cd^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cd(s)$

    $Cd^{2+}(aq) + Ba(s) \rightarrow Ba^{2+}(aq) + Cd(s)$

    $Cd^{2+}(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + Cd(s)$
11. Sample answer: copper(II) nitrate, platinum(II) nitrate, and mercury(II) nitrate.
12. $Cu^{2+}(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + Cu(s)$

    $Pt^{2+}(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + Pt(s)$

    $Hg^{2+}(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + Hg(\ell)$

(Student textbook page 590)

13. $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

    $Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$
14. $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

    $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
15. $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$

    $Ag^+(aq) + e^- \rightarrow Ag(s)$
16. $Sn(s) + Pb^{2+}(aq) \rightarrow Sn^{2+}(aq) + Pb(s)$

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

    $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$
17. $Au^{3+}(aq) + 3Ag(s) \rightarrow 3Ag^+(aq) + Au(s)$

    $Ag(s) \rightarrow Ag^+(aq) + e^-$

    $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$
18. $3Zn(s) + 2Fe^{3+}(aq) \rightarrow 3Zn^{2+}(aq) + 2Fe(s)$

    $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

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

(Student textbook page 611)

19. Oxidation numbers are usually (but not always) the same as the charge that an atom in a compound would have if the electrons were completely held by the atom with the greatest electronegativity instead of being shared. Unlike the numbers used to indicate charge, oxidation numbers do not represent a charge. Also, the plus/minus signs for charges are written as a superscript after the number, whereas the plus/minus signs for oxidation numbers are written before the number. Additionally, oxidation numbers can be non-integers, whereas the numbers used to indicate the charge are not.
20. The oxidation number of any pure element is zero.
21. An oxidation number may be a fraction if an atom or ion with two or more different oxidation numbers, such as iron(III) and iron(II), is present in a compound. The fraction is the average of these oxidation numbers. Sometimes, the rules lead to fraction oxidation numbers in other cases as well. For instance, the oxidation numbers in a neutral molecule must add to zero. In acetone, $H_6C_3O(\ell)$, hydrogen has

an oxidation number of +1 and that of oxygen is −2. Thus, carbon must have an oxidation number of $-\frac{4}{3}$.

22. The sum of the oxidation numbers for the atoms in a neutral molecule must be zero.

23. The atom lost electrons or undergone oxidation.

24. In $SO_3(g)$, sulfur has an oxidation number of +6, and in $H_2SO_4(aq)$, the sulfur has an oxidation number of +6. With no change to the oxidation number, sulfur is neither oxidized nor reduced in this reaction.

Answers to Caption Questions

Figure 9.3 (Student textbook page 586): The blue colour is characteristic of $Cu^{2+}(aq)$. In the second test tube, some of the $Cu^{2+}(aq)$ has been reduced to Cu(s), so there is less of this ion in solution and the blue colour has faded. In the third test tube, all of the $Cu^{2+}(aq)$ has been reduced and there is no colour.

Figure 9.8 (Student textbook page 600): The reaction $2C(s) + O_2(g) \rightarrow 2CO(g)$ is exothermic.

Figure 9.11 (Student textbook page 603): When chemical bonds are broken and re-formed, energy, in the form of visible light, is released.

Answers to Section 9.1 Review Questions

(Student textbook page 589)

1. Oxidation is a loss of electrons by a substance in a reaction, while a reduction is a gain in electrons by a substance in a reaction. Since electrons do not exist freely in solution, the electrons lost must be gained, and so reduction and oxidations always occur together.

2. Elements such as halogens or sulfur (elements with a relatively high electronegativity will gain electrons, thus being reduced, which by definition means that they cause an oxidation).

3. The oxidizing agent causes another substance to go through oxidation, meaning that this other substance has lost electrons. The material that causes the loss of electrons will be the material that gains these electrons, and by definition, this means the material has undergone reduction.

4. Lithium will act as the reducing agent, as it will go from metallic lithium to lithium ions by losing an electron. This loss of an electron means that lithium has gone through oxidation, which is true of a reducing agent in a redox reaction.

5. **a.** magnesium is the reducing agent and the hydrogen ion is the oxidizing agent

   b. sodium is the reducing agent and hydrogen ion is the oxidizing agent

   c. iron(II) ion is the oxidizing agent and chromium is the reducing agent

6. Possible answers for each part are shown below. Equations must be balanced. Oxidizing agents must gain electrons, and reducing agents must lose electrons. Oxidizing agents react with metals higher in the activity series; reducing agents react with metals lower in the activity series.

   a. $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$

   b. $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$

   c. $Al(s) + 3Au^{+}(aq) \rightarrow Al^{3+}(aq) + 3Au(s)$

   d. $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

7. **a.** $Al(s) + Fe^{3+}(aq) \rightarrow Fe(s) + Al^{3+}(aq)$

   b. Al(s) is oxidized and $Fe^{3+}(aq)$ is reduced

   c. $Fe^{3+}(aq)$ is the oxidizing agent and Al(s) is the reducing agent

8. **a.** $CuSO_4(aq) + Ca(s) \rightarrow Cu(s) + CaSO_4(aq)$

   $Cu^{2+}(aq) + \cancel{SO_4^{2-}(aq)} \rightarrow Cu(s) + Ca^{2+}(aq) + \cancel{SO_4^{2-}(aq)}$

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

   Ca(s) loses $2e^-$; $Cu^{2+}(aq)$ gains $2e^-$

   b. $Hg(NO_3)_2(aq) + Ni(s) \rightarrow Hg(\ell) + Ni(NO_3)_2(aq)$

   $Hg^{2+}(aq) + \cancel{2NO_3^{-}(aq)} + Ni(s) \rightarrow Hg(\ell) + Ni^{2+}(aq) + \cancel{2NO_3^{-}(aq)}$

   $Hg^{2+} + Ni(s) \rightarrow Ni^{2+}(aq) + Hg(\ell)$

   Ni(s) loses $2e^-$, $Hg^{2+}(aq)$ gains $2e^-$

   c. Note that AgCl and TlCl are compounds of low solubility and this reaction will occur only to a limited extent.

   $AgCl(aq) + Tl(s) \rightarrow Ag(s) + TlCl(aq)$

   $Ag^{+}(aq) + \cancel{Cl^{-}(aq)}\ Tl(s) \rightarrow Ag(s) + Tl^{+}(aq) + \cancel{Cl^{-}(aq)}$

   $Ag^{+}(aq) + Tl(s) \rightarrow Tl^{+}(aq) + Ag(s)$

   Tl(s) loses $1e^-$, $Ag^{+}(aq)$ gains $1e^-$

9. Elements such as iron, zinc or magnesium (any element that is a stronger reducing agent than cobalt) can be used to cause a spontaneous reaction.

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

   $Zn(s) + Co^{2+}(aq) \rightarrow Co(s) + Zn^{2+}(aq)$

   $Mg(s) + Co^{2+}(aq) \rightarrow Co(s) + Mg^{2+}(aq)$

10. Combustion of a hydrocarbon e.g.

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Carbon changes from an oxidation number of −4 in $CH_4(g)$ to +4 in $CO_2(g)$. This loss of electrons is an oxidation. Oxygen changes from an oxidation number of 0 in $O_2(g)$ to −2 in $H_2O(g)$. This gain of electrons is a reduction.

11. It is not possible because you cannot have two oxidation reactions and no reduction reaction.

12. These photos indicate that a redox reaction had spontaneously occurred, to show that the metal is a stronger reducing agent than silver.

Answers to Section 9.2 Review Questions

(Student textbook page 602)

1. a. *Oxidation*: $2I^-(aq) \rightarrow I_2(s) + 2e^-$

Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

Balanced redox reaction:

$Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(s)$

b. *Oxidation*: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

Reduction: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$

Balanced redox reaction:

$2Ag^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + 2Ag(s)$

c. *Oxidation*: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$

Reduction: $Al^{3+}(aq) + 3e^- \rightarrow Al(s)$

Balanced redox reaction:

$2Al^{3+}(aq) + 3Ni(s) \rightarrow 3Ni^{2+}(aq) + 2Al(s)$

2. Since the nitrate ions are spectator ions and do not undergo reduction or oxidation, they can be removed from the equation and added in at the end.

3. The labelled flowchart should show the following steps:

1. How many oxygen molecules in the metal oxide? (Add this many water molecules to the left side of the half reaction.)

2. How many hydrogen atoms are now on the left side? (This will be double the number of oxygen molecules just added.)

3. How many $H^+(aq)$ are needed on the right side to balance hydrogen? (These will be added to the right side to balance the hydrogen.)

4. How many hydroxide ions must be added to convert hydrogen ions to water? (Add this number of hydroxide ions to both sides of the half reaction, and convert the hydrogen ions and hydroxide ions to water molecules, and reduce the water molecules that now exist on both sides of the half reaction.)

5. What is the charge on each side of the half reaction, and which side requires electrons to be added to balance the charge? (Add this number of electrons to the appropriate side to balance the charge.)

4. a. $14H^+(aq) + Cr_2O_7^{2-}(aq) + 6e^- \rightarrow$
$2Cr^{3+}(aq) + 7H_2O(\ell)$; reduction

b. $S_2O_3^{2-}(aq) + 3H_2O(\ell) \rightarrow$
$S_2O_6^{2-}(aq) + 6H^+(aq) + 6e^-$; oxidation

c. $20H^+(aq) + 4AsO_4^{3-}(aq) + 8e^- \rightarrow$
$As_4O_6(aq) + 10H_2O(\ell)$; reduction

d. $12OH^-(aq) + Br_2(g) \rightarrow$
$2BrO_3^-(aq) + 6H_2O(\ell) + 10e^-$; oxidation

5. The sequence of steps must be followed because the oxygen atoms are introduced in H_2O, which cannot be predicted by inspection.

6. a. *oxidation*: $Co(s) \rightarrow Co^{3+} + 3e^-$

reduction: $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$

$Au^{3+}(aq) + Co(s) \rightarrow Au(s) + Co^{3+}(aq)$

b. *oxidation*: $Cu(s) \rightarrow Cu^{2+}(aq)$

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

reduction: $NO_3^-(aq) \rightarrow NO(g)$

$NO_3^-(aq) \rightarrow NO(g) + 2H_2O(\ell)$

$NO_3^-(aq) + 4H^+ \rightarrow NO(g) + 2H_2O(\ell)$

$NO_3^-(aq) + 4H^+ + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

multiply oxidation rx. by 3:

$3[Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-]$

$3\ Cu(s) \rightarrow 3Cu^{2+}(aq) + 6e^-$

multiply reduction rx. by 2:

$2[NO_3^-(aq) + 4H^+ + 3e^- \rightarrow NO(g) + 2H_2O(\ell)]$

$2NO_3^-(aq) + 8H^+ + 6e^- \rightarrow 2NO(g) + 4H_2O(\ell)$

combine:

$3Cu(s) + 2NO_3^-(aq) + 8H^+ \cancel{+ 6e^-} \rightarrow$
$3Cu^{2+}(aq) + \cancel{6e^-} + 2NO(g) + 4H_2O(\ell)$

balance:

$3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow$
$3Cu^{2+}(aq) + 2NO(g) + 4H_2O(\ell)$

c. *oxidation*: $Al(s) \rightarrow AlO_2^-(aq)$

$Al(s) + 2H_2O(\ell) \rightarrow AlO_2^-(aq)$

$Al(s) + 2H_2O(\ell) \rightarrow$
$AlO_2^-(aq) + 4H^+(aq)$

$Al(s) + 2H_2O(\ell) + 4OH^-(aq) \rightarrow$
$AlO_2^-(aq) + 4H^+(aq) + 4OH^-(aq)$

$Al(s) + 2H_2O(\ell) + 4OH^-(aq) \rightarrow$
$AlO_2^-(aq) + 4H_2O(\ell)$

$Al(s) + 4OH^-(aq) \rightarrow$
$AlO_2^-(aq) + 2H_2O(\ell) + 3e^-$

reduction: $NO_3^- \rightarrow NH_3(g)$

$NO_3^- \rightarrow NH_3(g) + 3H_2O(\ell)$

$NO_3^- + 9H^+(aq) \rightarrow NH_3(g) + 3H_2O(\ell)$

$NO_3^- + 9H^+(aq) + 9OH^-(aq) \rightarrow$
$NH_3(g) + 3H_2O(\ell) + 9OH^-(aq)$

$NO_3^- + 9H_2O(\ell) \rightarrow$
$NH_3(g) + 3H_2O(\ell) + 9OH^-(aq)$

$NO_3^- + 6H_2O(\ell) + 8e^- \rightarrow$
$NH_3(g) + 9OH^-(aq)$

multiply oxidation rx. by 8:

$8[Al(s) + 4OH^-(aq) \rightarrow$
$AlO_2^-(aq) + 2H_2O(\ell) + 3e^-]$

$8Al(s) + 32OH^-(aq) \rightarrow$
$8AlO_2^-(aq) + 16H_2O(\ell) + 24e^-]$

multiply reduction rx. by 3:

$3[NO_3^- + 6H_2O(\ell) + 8e^- \rightarrow NH_3(g) + 9OH^-(aq)]$

$3NO_3^- + 18H_2O(\ell) + 24e^- \rightarrow$
$3NH_3(g) + 27OH^-(aq)$

combine:

$8Al(s) + 32OH^-(aq) + 3NO_3^- + 18H_2O(\ell) + \cancel{24e^-}$
$\rightarrow 8AlO_2^-(aq) + 16H_2O(\ell) + \cancel{24e^-} + 3NH_3(g)$
$+ 27OH^-(aq)$

balance:

$3NO_3^-(aq) + 8Al(s) + 2H_2O(\ell) + 5OH^-(aq) \rightarrow$
$3NH_3(g) + 8AlO_2^-(aq)$

7. *Possible answer*: a flow chart with labels:

1. How many oxygen molecules in the metal oxide? (Add this many water molecules to the right side of the half reaction.)

2. How many hydrogen atoms are now on the right side? (This will be double the number of oxygen molecules just added.)

3. How many $H^+(aq)$ are needed on the left side to balance hydrogen? (These will be added to the left side to balance the hydrogen.)

4. What is the charge on each side of the half reaction, and which side requires electrons to be added to balance the charge? (Add this number of electrons to the appropriate side to balance the charge.)

8. The process is called refining. It produces steel of higher purity that can be alloyed with other metals for specific uses

molten pig iron from blast furnace
(5% C + Si, P, S + other impurities)

↓

oxygen + lime (CaO)

↓

$SiO_2(s) + CaO(s) \rightarrow CaSiO_3(\ell)$

$P_4O_{10}(\ell) + 6CaO(s) \rightarrow 2Ca_3(PO_4)_2(\ell)$

↓

iron (higher purity)

9. *Half-reactions*:

$Hg^+(aq) + e^- \cancel{(aq)} \rightarrow Hg(\ell)$

$Hg^+(aq) \rightarrow Hg^{2+}\cancel{(aq)} + e^-(aq)$

Net ionic equation: $2Hg^+(aq) \rightarrow Hg(\ell) + Hg^{2+}(aq)$

10. a. *oxidation*: $I^-(aq) \rightarrow HIO(aq)$

$I^-(aq) + H_2O(\ell) \rightarrow HIO(aq)$

$I^-(aq) + H_2O(\ell) \rightarrow HIO(aq) + H^+(aq)$

$I^-(aq) + H_2O(\ell) \rightarrow$
$HIO(aq) + H^+(aq) + 2e^-$

reduction: $HClO_2(aq) \rightarrow Cl_2(g)$

$2HClO_2(aq) \rightarrow Cl_2(g)$

$2HClO_2(aq) \rightarrow Cl_2(g) + 4H_2O(\ell)$

$2HClO_2(aq) + 6H^+(aq) \rightarrow$
$Cl_2(g) + 4H_2O(\ell)$

$2HClO_2(aq) + 6H^+(aq) + 6e^- \rightarrow$
$Cl_2(g) + 4H_2O(\ell)$

multiply oxidation rx. by 3:

$3[I^-(aq) + H_2O(\ell) \rightarrow HIO(aq) + H^+(aq) + 2e^-]$

$3I^-(aq) + 3H_2O(\ell) \rightarrow 3HIO(aq) + 3H^+(aq) + 6e^-]$

multiply reduction rx. by 1:

$2HClO_2(aq) + 6H^+(aq) + 6e^- \rightarrow Cl_2(g) + 4H_2O(\ell)$

combine:

$2HClO_2(aq) + 6H^+(aq) + \cancel{6e^-} + 3I^-(aq) + 3H_2O(\ell)$
$\rightarrow 3HIO(aq) + 3H^+(aq) + \cancel{6e^-} + Cl_2(g) + 4H_2O(\ell)$

balance: $2HClO_2(aq) + 3I^-(aq) + 3H^+(aq) \rightarrow$
$Cl_2(g) + 3HIO_3(aq) + H_2O(\ell)$

b. *oxidation*: $P_4(s) \rightarrow 4H_2PO_4^-(aq)$

$P_4(s) + 16H_2O(\ell) \rightarrow 4H_2PO_4^-(aq)$

$P_4(s) + 16H_2O(\ell) \rightarrow$
$4H_2PO_4^-(aq) + 24H^+(aq)$

$P_4(s) + 16H_2O(\ell) \rightarrow$
$4H_2PO_4^-(aq) + 24H^+(aq) + 20e^-$

reduction: $IO_3^-(aq) \rightarrow I^-(aq)$

$IO_3^-(aq) \rightarrow I^-(aq) + 3H_2O(\ell)$

$IO_3^-(aq) + 6H^+(aq) \rightarrow$
$I^-(aq) + 3H_2O(\ell)$

$IO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow$
$I^-(aq) + 3H_2O(\ell)$

multiply oxidation rx. by 3:

$3[P_4(s) + 16H_2O(\ell) \rightarrow 4H_2PO_4^-(aq) + 24H^+(aq) + 20e^-]$

$3P_4(s) + 48H_2O(\ell) \rightarrow 12H_2PO_4^-(aq) + 72H^+(aq) + 60e^-]$

multiply reduction rx. by 10:

$10[IO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow I^-(aq) + 3H_2O(\ell)]$

$10IO_3^-(aq) + 60H^+(aq) + 60e^- \rightarrow 10\ I^-(aq) + 30H_2O(\ell)]$

combine:

$3P_4(s) + 48H_2O(\ell) + 10IO_3^-(aq) + 60H^+(aq) + \cancel{60e^-} \rightarrow 12H_2PO_4^-(aq) + 72H^+(aq) + \cancel{60e^-} + 10\ I^-(aq) + 30H_2O(\ell)$

balance:

$3P_4(s) + 10IO_3^-(aq) + 18H_2O(\ell) \rightarrow 12H_2PO_4^-(aq) + 10I^- + 12H^+(aq)$

c. *oxidation*: $Hg(\ell) \rightarrow Hg^{2+}(aq)$

$Hg(\ell) \rightarrow Hg^{2+}(aq) + 2e^-$

reduction: $Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$

$Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq)$

$Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$

$Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$

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

multiply oxidation rx. by 3:

$3[Hg(\ell) \rightarrow Hg^{2+}(aq) + 2e^-]$

$3Hg(\ell) \rightarrow 3Hg^{2+}(aq) + 6e^-$

multiply reduction rx. by 1:

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

combine:

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

balance:

$Cr_2O_7^{2-}(aq) + 3Hg(\ell) + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3Hg^{2+}(aq) + 7H_2O(\ell)$

d. *oxidation*: $H_2O_2(aq) \rightarrow O_2(g)$

$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq)$

$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$

reduction: $MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$

$MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

$MnO_4^-(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

multiply oxidation rx. by 5:

$5[H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-]$

$5H_2O_2(aq) \rightarrow 5O_2(g) + 10H^+(aq) + 10e^-$

multiply reduction rx. by 2:

$2[MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)]$

$2MnO_4^-(aq) + 16H^+(aq) + 10e^- \rightarrow 2Mn^{2+}(aq) + 8H_2O(\ell)$

combine:

$5H_2O_2(aq) + 2MnO_4^-(aq) + 16H^+(aq) + \cancel{10e^-} \rightarrow 5O_2(g) + 10H^+(aq) + \cancel{10e^-} + 2Mn^{2+}(aq) + 8H_2O(\ell)$

balance:

$2MnO_4^-(aq) + 5H_2O_2(\ell) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 8H_2O(\ell)$

11. Add the same number of hydroxide ions to both sides as there hydrogen ions in the balanced acidic reaction. The side with the hydrogen ions will form water molecules with the hydrogen ions and hydroxide ions, which will then allow for the reduction of the number of water molecules in the overall reaction.

12. a. $2HClO_2(aq) + 3I^-(aq) \rightarrow 3Cl_2(g) + 3HIO_3(aq) + 3OH^-(aq)$

b. $3P_4(s) + 10IO_3^-(aq) + 12OH^- \rightarrow 12H_2PO_4^-(aq) + 10I^-(aq) + 6H_2O(\ell)$

c. $Cr_2O_7^{2-}(aq) + 3Hg(\ell) + 7H_2O(\ell) \rightarrow 2Cr^{3+}(aq) + 3Hg^{2+}(aq) + 14OH^-(aq)$

d. $2MNO_4^-(aq) + 5H_2O_2(\ell) + 5H_2O(\ell) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 6OH^-(aq)$

13. a. *oxidation*: $Br^-(aq) \rightarrow BrO_3^-(aq)$

$Br^-(aq) + 3H_2O(\ell) \rightarrow BrO_3^-(aq)$

$Br^-(aq) + 3H_2O(\ell) \rightarrow BrO_3^-(aq) + 6H^+(aq)$

$Br^-(aq) + 3H_2O(\ell) \rightarrow BrO_3^-(aq) + 6H^+(aq) + 6e^-$

reduction: $ClO_4^-(aq) \rightarrow Cl^-(aq)$

$ClO_4^-(aq) \rightarrow Cl^-(aq) + 4H_2O(\ell)$

$ClO_4^-(aq) + 8H^+(aq) \rightarrow Cl^-(aq) + 4H_2O(\ell)$

$ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightarrow Cl^-(aq) + 4H_2O(\ell)$

multiply oxidation rx. by 4:

$4[Br^-(aq) + 3H_2O(\ell) \rightarrow BrO_3^-(aq) + 6H^+(aq) + 6e^-]$

$4Br^-(aq) + 12H_2O(\ell) \rightarrow 4BrO_3^-(aq) + 24H^+(aq) + 24e^-$

multiply reduction rx. by 3:
$3[ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightarrow Cl^-(aq) + 4H_2O(\ell)]$
$3ClO_4^-(aq) + 24H^+(aq) + 24e^- \rightarrow 3Cl^-(aq) + 12H_2O(\ell)$
combine:
$4Br^-(aq) + \cancel{12H_2O(\ell)} + 3ClO_4^-(aq) + \cancel{24H^+(aq)} + \cancel{24e^-} \rightarrow 4BrO_3^-(aq) + \cancel{24H^+(aq)} + \cancel{24e^-} + 3Cl^-(aq) + \cancel{12H_2O(\ell)}$
balance:
$3ClO_4^-(aq) + 4Br^-(aq) \rightarrow 3Cl^-(aq) + 4BrO_3^-(aq)$

b. *oxidation*:
$N_2O_4(g) \rightarrow NO_3^-(aq)$
$N_2O_4(g) \rightarrow 2NO_3^-(aq)$
$N_2O_4(g) + 2H_2O(\ell) \rightarrow 2NO_3^-(aq)$
$N_2O_4(g) + 2H_2O(\ell) \rightarrow 2NO_3^-(aq) + 4H^+(aq)$
$N_2O_4(g) + 2H_2O(\ell) + 4OH^-(aq) \rightarrow 2NO_3^-(aq) + 4H^+(aq) + 4OH^-(aq)$
$N_2O_4(g) + 2H_2O(\ell) + 4OH^-(aq) \rightarrow 2NO_3^-(aq) + 4H_2O(\ell)$
$N_2O_4(g) + 2H_2O(\ell) + 4OH^-(aq) \rightarrow 2NO_3^-(aq) + 4H_2O(\ell) + 2e^-$
reduction:
$TeO_3^{2-}(aq) \rightarrow Te(s)$
$TeO_3^{2-}(aq) \rightarrow Te(s) + 3H_2O(\ell)$
$TeO_3^{2-}(aq) + 6H^+(aq) + 6OH^-(aq) \rightarrow Te(s) + 3H_2O(\ell) + 6OH^-(aq)$
$TeO_3^{2-}(aq) + 6H_2O(\ell) \rightarrow Te(s) + 3H_2O(\ell) + 6OH^-(aq)$
$TeO_3^{2-}(aq) + 3H_2O(\ell) \rightarrow Te(s) + 6OH^-(aq)$
$TeO_3^{2-}(aq) + 3H_2O(\ell) + 4e^- \rightarrow Te(s) + 6OH^-(aq)$
multiply oxidation rx. by 2:
$2[N_2O_4(g) + 2H_2O(\ell) + 4OH^-(aq) \rightarrow 2NO_3^-(aq) + 4H_2O(\ell) + 2e^-]$
$2N_2O_4(g) + 4H_2O(\ell) + 8OH^-(aq) \rightarrow 4NO_3^-(aq) + 8H_2O(\ell) + 4e^-$
multiply reduction rx. by 1:
$TeO_3^{2-}(aq) + 6H_2O(\ell) + 4e^- \rightarrow Te(s) + 3H_2O(\ell) + 6OH^-(aq)$
combine:
$2N_2O_4(g) + 4H_2O(\ell) + 8OH^-(aq) + TeO_3^{2-}(aq) + 3H_2O(\ell) + \cancel{4e^-} \rightarrow 4NO_3^-(aq) + 8H_2O(\ell) + \cancel{4e^-} + Te(s) + 6OH^-(aq)$
balance:
$2N_2O_4(g) + TeO_3^{2-}(aq) + 2OH^-(aq) \rightarrow Te(s) + 4NO_3^-(aq) + H_2O(\ell)$

c. *oxidation*: $H_2O_2(aq) \rightarrow O_2(g)$
$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq)$
$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$
reduction: $Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$
$Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq)$
$Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$
multiply oxidation rx. by 3:
$3[H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-]$
$3H_2O_2(aq) \rightarrow 3O_2(g) + 6H^+(aq) + 6e^-$
multiply reduction rx. by 1:
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$
combine:
$3H_2O_2(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) + \cancel{6e^-} \rightarrow 3O_2(g) + 6H^+(aq) + \cancel{6e^-} + 2Cr^{3+}(aq) + 7H_2O(\ell)$
balance: $3H_2O_2(\ell) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3O_2(g) + 2Cr^{3+}(aq) + 7H_2O(\ell)$

14. It is a redox reaction because sulfur is both being oxidized from an oxidation number of −2 in $H_2S(g)$ to an oxidation number of 0 in S(s) and reduced from an oxidation number of +4 in $SO_2(g)$ to an oxidation number of 0 in S(s).

Answers to Section 9.3 Review Questions

(Student textbook page 617)

1. Fluorine is the most electronegative element in the periodic table. Therefore, it does not "lose" electrons in chemical bonds. It has seven valence electrons, and thus forms bonds in which it gains one valence electron, satisfying the octet rule. Because it is in the second row of the periodic table, its outer occupied energy level cannot accommodate more than eight electrons (as bromine, iodine, or chlorine can). Therefore, it does not form bonds in which it gains more than one electron, and its oxidation number is always −1.

2. a. MnO_2 +4
b. +7
c. + 6
d. $KMnO_4$ +7

3. Zero in the diatomic molecule, $O_2(g)$; −1 in peroxides, $H_2O_2(aq)$; −2 in most anions, $CO_3^{2-}(aq)$; +2 in an oxygen difluoride, $OF_2(g)$

4. a. yes
b. no
c. no
d. yes

5. Students may suggest definitions similar to the following:

- A redox reaction is a chemical reaction in which there is a change in the oxidation number of some or all of the atoms of the reactants.
- A redox reaction is a chemical reaction in which atoms of one element are oxidized and atoms of another element are reduced.
- A redox reaction is a chemical reaction in which electrons are transferred among reactants. The number of electrons gained by the oxidizing agent must equal the number of electrons lost by the reducing agent.

6. When one element combines with another element in a chemical reaction, the reaction is always a redox reaction. The atoms in uncombined elements have an oxidation number of 0. However, different elements have different electronegativities. Therefore, when they combine in a compound, atoms of the less electronegative element are oxidized and atoms of the more electronegative element are reduced.

7. The total increase in oxidation numbers corresponds to the number of electrons lost by oxidation. The total decrease in oxidation numbers corresponds to the number of electrons gained by reduction. In a reaction, net gain of electrons must be equal to net loss of electrons. Thus, the total increase in oxidation numbers must equal the total decrease in oxidation numbers.

8. $4NH_3(g) + 7O_2(g) \rightarrow 4NO_2(g) + 6H_2O(\ell)$

9. a. $+2$

b. Using a Lewis structure, students may arrive at two different oxidation numbers for sulfur using the following reasoning. The central sulfur atom usually has 6 electrons in its valence shell. The three oxygen atoms, however, have a greater electronegativity than sulfur. Therefore, they can be considered to have taken these electrons, leaving the central sulfur with a $+6$ oxidation number. The sulfur atom bonded to the central sulfur atom may be considered to share its electrons with the central sulfur. Therefore, it will have an oxidation number of -1. Since the central sulfur $(+6)$ can now be considered to have one shared electron it will have a total oxidation number of $+6 + (-1) = +5$.

c. Using the first method, the oxidation number for both sulfur atoms is $+2$. Using the second method, the oxidation number for one sulfur atom is -1, and for the other is $+5$. (Notice the average of the second result gives the first result).

d. Using Lewis structures allows for the fact that atoms of the same element in a molecule or polyatomic ion may have different oxidation numbers, depending on the bonding arrangement. The disadvantage to this method is that it is more difficult and time-consuming than using the rules.

e. The advantage to using the oxidation number rules is that they apply to most cases and are quick and easy to use. The disadvantage to using these rules is that they do not allow for the possibility of different oxidation numbers for the same atoms in a molecule or polyatomic ion.

10. a. $CH_3COOH(aq) + O_2(g) \rightarrow CO_2(g) + H_2O(\ell)$

Oxidation number method:

$$\underset{0\ +1\ \ 0\ -2\ -2\ +1}{CH_3COOH(aq)} + \underset{0}{O_2(g)} \rightarrow \underset{+4\ -2}{CO_2(g)} + \underset{+1\ \ -2}{H_2O(\ell)}$$

1(-4e⁻) = -4e⁻

$$CH_3COOH(aq) + O_2(g) \rightarrow CO_2(g) + H_2O(\ell)$$

2(+2e⁻) = +4e⁻

(Note: The zero oxidation numbers for the carbon atoms in ethanoic acid represent an average value.)

The coefficient 1 goes in front of the ethanoic acid and the coefficient 2 goes in front of the oxygen.

$CH_3COOH(aq) + 2O_2(g) \rightarrow CO_2(g) + H_2O(\ell)$

Balance the right side of the equation by inspection, ensuring that the ratios of the carbon and oxygen atoms remain as above:

$CH_3COOH(aq) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(\ell)$

b. $O_2(g) + H_2SO_3(aq) \rightarrow HSO_4^-\ (aq)$

Because the reaction takes place in acidic conditions, add H^+ to balance the charge. Then balance by inspection.

$O_2(g) + 2H_2SO_3(aq) \rightarrow 2HSO_4^-(aq) + 2H^+(aq)$

11. a. $NH_3(g) + Cl_2(g) \rightarrow NH_4Cl(aq) + N_2(g)$

Assign oxidation numbers.

$$\underset{-3\ +1}{NH_3(g)} + \underset{0}{Cl_2(g)} \rightarrow \underset{-3\ +1\ \ -1}{NH_4Cl(aq)} + \underset{0}{N_2(g)}$$

Notice among the products, that there are two nitrogen atoms with different oxidation numbers, -3 and 0. Some of the nitrogen atoms in the reactants will not change oxidation numbers, while others will change from -3 to 0. Therefore, to keep things clear, create two sets of ammonia molecules on the reactant side.

$$\underset{-3\ +1}{NH_3(g)} + \underset{-3\ +1}{NH_3(g)} + \underset{0}{Cl_2(g)} \rightarrow \underset{-3\ +1\ \ -1}{NH_4Cl(aq)} + \underset{0}{N_2(g)}$$

Now track the gain and loss of electrons. There is only one nitrogen atom in NH_3, but two chlorine atoms in Cl_2. When the ratio of the number of atoms undergoing oxidation and reduction differs, you must account for all of the electrons gained or lost by the molecule. So, when each atom in Cl_2 gains 1 electron, the molecule gains 2 electrons.

$2(-3e^-) = -6e^-$

$$NH_3(g) + NH_3(g) + Cl_2(g) \rightarrow NH_4Cl(aq) + N_2(g)$$

$3(+2e^-) = +6e^-$

Place a coefficient of 2 in front of the first ammonia symbol and a coefficient of 3 in front of the chlorine symbol.

$2NH_3(g) + NH_3(g) + 3Cl_2(g) \rightarrow NH_4Cl(aq) + N_2(g)$

The ratio of these coefficients cannot change. Balance the equation by adding coefficients in front of the other compounds where needed.

$2NH_3(g) + 6NH_3(g) + 3Cl_2(g) \rightarrow 6NH_4Cl(aq) + N_2(g)$

Now that the equation is balanced, combine the two ammonia compounds among the reactants.

$8NH_3(g) + 3Cl_2(g) \rightarrow 6NH_4Cl(aq) + N_2(g)$

b. $Mn_3O_4(aq) + Al(s) \rightarrow Al_2O_3(aq) + Mn(s)$

Assign oxidation numbers.

$$\underset{+\frac{8}{3}\ \ -2}{Mn_3O_4}(aq) + \underset{0}{Al}(g) \rightarrow \underset{+3\ \ -2}{Al_2O_3}(aq) + \underset{0}{Mn}(s)$$

The oxidation number of +8/3 is an average of the three manganese atoms. When comparing the number of electrons gained and lost, you must consider the ratio of atoms. In the reactants there are three manganese atoms and one aluminum atom. Therefore you must consider the total number of electrons gained or lost for all of the atoms in the compound. So, when the three manganese atoms are reduced from an oxidation number of +8/3, there are a total of 8 electrons gained. When one aluminum atom is oxidized form an oxidation number of +3 to 0, it loses 3 electrons.

$3(+8e^-) = +24e^-$

$$Mn_3O_4(aq) + Al(g) \rightarrow Al_2O_3(aq) + Mn(s)$$

$8(-3e^-) = -24e^-$

Place the coefficient 8 in front of the Al and the coefficient 3 in front of the Mn_3O_4.

$3Mn_3O_4(aq) + 8Al(s) \rightarrow Al_2O_3(aq) + Mn(s)$

Balance the rest of the equation by inspection.

$3Mn_3O_4(aq) + 8Al(s) \rightarrow 4Al_2O_3(aq) + 9Mn(s)$

12. Oxidation numbers are values assigned to elements in a redox reaction to keep track of changes in charge. This system of book keeping is developed on a base of assumptions, that may seem arbitrary at times, but in fact do allow for the tracking of changes in charge.

Answers to Practice Problems

For full solutions to Practice Problems, see Part B of this Solutions Manual.

(Student textbook page 598)

1. $MnO_4^-(aq) + 5Ag(s) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Ag^+(aq) + 4H_2O(\ell)$

Oxidizing agent: MnO_4^- (aq); reducing agent: Ag(s)

2. $Hg(\ell) + 2NO_3^-(aq) + 4Cl^-(aq) + 4H^+(aq) \rightarrow HgCl_4^{2-}(s) + 2NO_2(g) + 2H_2O(\ell)$

Oxidizing agent: NO_3^- (aq); reducing agent: $Hg(\ell)$

3. $4H_2O(\ell) + AsH_3(s) + 4Zn^{2+}(aq) \rightarrow H_3AsO_4(aq) + 4Zn(s) + 8H^+(aq)$

Oxidizing agent: Zn^{2+}(aq); reducing agent: AsH_3(s)

4. $I_2(s) + 5ClO^-(aq) + H_2O(\ell) \rightarrow 2IO_3^-(aq) + 5Cl^-(aq) + 2H^+(aq)$

Oxidizing agent: ClO^-(aq); reducing agent I_2(s)

5. $6MnO_4^-(aq) + I^-(aq) + 6OH^-(aq) \rightarrow 6MnO_4^{2-}(aq) + IO_3^-(aq) + 3H_2O(\ell)$

Oxidizing agent: MnO_4^-(aq); reducing agent: I^-(aq)

6. $3H_2O_2(aq) + 2ClO_2(aq) + 2OH^-(aq) \rightarrow 2ClO^-(aq) + 3O_2(g) + 4H_2O(\ell)$

Oxidizing agent: ClO_2(aq); reducing agent: H_2O_2(aq)

7. $6ClO^-(aq) + 2CrO_2^-(aq) + 2H_2O(\ell) \rightarrow 2CrO_4^{2-}(aq) + 3Cl_2(g) + 4OH^-(aq)$

oxidizing agent: ClO^-(aq);
reducing agent: CrO_2^-(aq)

8. $4Al(s) + 3NO^-(aq) + 4H_2O(\ell) + OH^-(aq) \rightarrow 3NH_3(g) + 4AlO_2^-(aq)$

Oxidizing agent: NO^-(aq); reducing agent: Al(s)

9. $ClO_3^-(aq) + 2MnO_2(s) + 2OH^-(aq) \rightarrow Cl^-(aq) + 2MnO_4^-(aq) + H_2O(\ell)$

Oxidizing agent: ClO_3^- (aq); reducing agent: MnO_2(s)

10. $5PbO_2(s) + I_2(s) + 8H^+(aq) \rightarrow 5Pb^{2+}(aq) + 2IO_3^-(aq) + 4H_2O(\ell)$

Oxidizing agent: PbO_2(s); reducing agent: I_2(s)

(Student textbook page 606)

11. +3

12. 0

13. +6

14. +5

15. 0

16. -1

17. H $+1$, S $+4$, O -2

18. O -2, H $+1$

19. H $+1$, P $+5$, O -2

20. 0

21. -1

22. Al $+3$, H $+1$, C $+4$, O -2

23. N -3, H $+1$, P $+5$, O -2

24. K $+1$, H $+1$, I $+7$, O -2

(Student textbook page 611)

25. redox reaction

26. not a redox reaction

27. redox reaction

28. disproportionation redox reaction

29. redox reaction

30. not a redox reaction

31. redox reaction

32. (25) reducing agent: $H_2(g)$, oxidizing agent: $O_2(g)$

(27) reducing agent: $C_2H_6(g)$, oxidizing agent: $O_2(g)$

(28) $NO_2(g)$ is both the oxidizing and reducing agent.

(29) reducing agent: HI(aq), oxidizing agent:

$HMnO_4(aq)$ (31) reducing agent: $H_2(g)$, oxidizing agent: $NO_2(g)$

33. Bromine liquid is reduced and $ClO_2^-(aq)$ is oxidized.

34. The oxidation number of nickel decreases from $+2$ in the sulfide to 0 in its elemental form. This is a reduction. A reducing agent must be used to achieve this change, so this is a redox reaction. The same analysis can be used with copper sulfide and copper metal, with the same conclusion.

(Student textbook page 615)

35. $CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$

36. $B_2O_3(aq) + 6Mg(s) \rightarrow 3MgO(s) + Mg_3B_2(aq)$

37. $8H_2S(g) + 8H_2O_2(aq) \rightarrow S_8(s) + 16H_2O(\ell)$

38. $2OH^-(aq) + 2H_2O(\ell) + 2ClO_2(aq) + SbO_2^-(aq) \rightarrow 2ClO_2^-(aq) + Sb(OH)_6^-(aq)$

39. $14H^+(aq) + Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$

40. $I_2(g) + 10NO_3^-(aq) + 8H^+(aq) \rightarrow 2IO_3^-(aq) + 10NO_2(g) + 4H_2O(\ell)$

41. $2PbSO_4(aq) + 2H_2O(\ell) \rightarrow Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq) + 4H^+(aq)$

42. $H_2O(\ell) + 3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq) + 2\ OH^-(aq)$

43. $H_2O(\ell) + 3Ni(s) + 2MnO_4^-(aq) \rightarrow 3NiO(s) + 2MnO_2(s) + 2OH^-(aq)$

44. $6OH^-(aq) + I^-(aq) + 6Ce^{4+}(aq) \rightarrow IO_3^-(aq) + 6Ce^{3+}(aq) + 3H_2O(\ell)$

Answers to Chapter 9 Review Questions

(Student textbook pages 624-9)

1. e

2. a

3. b

4. b

5. c

6. c

7. e

8. a

9. c

10. a

11. e

12. a

13. c

14. c

15. Methane is being oxidized, and oxygen is being reduced.

16. $H_2(g)$ is being oxidized; $O_2(g)$ is being reduced.

17. The reducing agent causes the reduction, which means that it causes another reactant to gain electrons that it loses. Since it loses the electrons, this material by definition is going through oxidation.

18. A is being oxidized as it goes through the half reaction: $A \rightarrow A^+ + e^-$, and B is being reduced as it goes through the half reaction $B^+ + e^- \rightarrow B$

19. Aluminum is the reducing agent, and the iodine is the oxidizing agent.

20. a. $3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$

b. *Ionic equation*:

$3Cu(s) + 8NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 6NO_3^-(aq) + 2NO(g) + 4H_2O(\ell)$

Net ionic equation:

$3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(\ell)$

c. *Oxidizing agent*: $HNO_3(aq)$; reducing agent: Cu(s)

d. *Oxidation*: $Cu(s) \rightarrow Cu^{+2}(aq) + 2e^-$

Reduction: $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

e. 0 to +2

21. a. metallic element

b. non-metallic element

c. non-metallic element

d. metallic element

22. a. Ba +2, Cl −1

b. Al +3, C −4

c. Ag +1, S +6, O −2

d. N +4, O -2

e. N −3, H +1, C +3, O −2

f. S 0

g. As +3, O −2

h. V +4, O −2

i. Xe +4, F −1

j. S $+\frac{9}{4}$, O −2

23. a. reduction

b. oxidation

c. reduction

d. oxidation

e. oxidation

24. a. $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$

b. $CaO(s) + H_2O(\ell) \rightarrow Ca(OH)_2(aq)$

c. $2C_7H_5(NO_2)_3(s) \rightarrow 12CO(g) + 2C(s) + 5H_2(g) + 3N_2(g)$

d. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

25. Phosphorus is both reduced and oxidized. Phosphorus atoms, with oxidation number 0, are reduced in $PH_3(s)$, with oxidation number −3. They are oxidized in $H_3PO_4(aq)$, with oxidation number +5.

26. $I^-(aq)$

The oxidation number of iodine in I^- (aq) is −1 and in $I_2(aq)$ is 0. This represents a loss of electrons or an oxidation. In a redox reaction it is the reducing agent that undergoes oxidation.

27. a. $Li_3N(s)$ the oxidation number is −3

b. $HNO_3(aq)$ the oxidation number is +5

28. The assignment of oxidation numbers to the atoms will help identify which species in the reaction are undergoing a loss or gain of electrons. From this information, two half reactions can be written and the rules for balancing these half reactions can be followed to obtain a balanced equation.

29. a. $H_2O_2(aq) + 2Fe(OH)_2(s) \rightarrow 2Fe(OH)_3(s)$; a redox reaction

b. $PCl_3(\ell) + 3H_2O(\ell) \rightarrow H_3PO_3(aq) + 3HCl(aq)$; not a redox reaction

c. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(\ell)$; a redox reaction

d. $3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$; a redox and disproportionation reaction

e. $MnO_2(s) + 2Cl^-(aq) + 4H^+(aq) \rightarrow Mn^{2+}(aq) + Cl_2(g) + 2H_2O(\ell)$; a redox reaction

30. a. Oxidation: $Pb^{2+}(aq) \rightarrow Pb^{4+}(aq) + 2e^-$

Reduction: $Al^{3+}(aq) + 3e^- \rightarrow Al(s)$

b. Balanced equation: $3Pb^{2+}(aq) + 2Al^{3+}(aq) \rightarrow 3Pb^{4+}(aq) + 2Al(s)$;

$3Pb^{2+}(aq)$ ions and $2Al^{3+}(aq)$ are required.

31.

Strongest oxidizing agent	Weakest reducing agent
$X^{2+}(aq)$	X(s)
$Y^{2+}(aq)$	Y(s)
$W^{2+}(aq)$	W(s)
$Z^{2+}(aq)$	Z(s)
Weakest oxidizing agent	Strongest reducing agent

32. a. $O_2(g)$ is the oxidizing agent, and $C_6H_6(\ell)$ is the reducing agent.

b. The reaction is not a redox reaction because the oxidation numbers do not change.

c. $I_2(s)$ is the oxidizing agent, and $H_2(g)$ is the reducing agent.

d. $KMnO_4(aq)$ is the oxidizing agent, and CuCl(s) is the reducing agent.

e. Cu(s) is the reducing agent, and $Ag^+(aq)$ is the oxidizing agent.

f. The reaction is not a redox reaction because the oxidation numbers do not change.

g. $Mn^{2+}(aq)$ is the reducing agent, and $BiO_3^-(aq)$ is the oxidizing agent.

33. a. oxidation number +5 in $V_2O_5(s)$, $VO_2^+(aq)$ VO_3^- (aq), $VO_4^{3-}(aq)$, and $V_3O_9^{3-}(aq)$; oxidation number +4 in $VO_2(s)$ and $VO^{2+}(aq)$

b. This is not a redox reaction since there are no species that change oxidation number.

34. The chlorite ion, ClO_2^-, is an example of an ion in which chlorine has an oxidation number of +3.

The sum of the oxidation numbers of Cl and O must be −1.

$(Cl) + 2(-2) = -1$

$Cl = +3$

35. Oxygen is consumed and carbon dioxide is produced. Therefore, some of the oxygen taken in by the body is reduced, indicating a redox reaction must have taken place.

36. a. *oxidation*: $Cd(s) \rightarrow Cd^{2+}(aq)$

$Cd(s) + 2e^- \rightarrow Cd^{2+}(aq)$

reduction: $Co^{3+}(aq) \rightarrow Co^{2+}(aq)$

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

multiply oxidation rx. by 1: $Cd(s) + 2e^- \rightarrow Cd^{2+}(aq)$

multiply reduction rx. by 2:

$2[Co^{3+}(aq) \rightarrow Co^{2+}(aq) + e^-]$

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

combine: $Cd(s) + \cancel{2e^-} + 2Co^{3+}(aq) \rightarrow Cd^{2+}(aq) + 2Co^{2+}(aq) + \cancel{2e^-}$

balance and add spectator ion:

$Cd(s) + 2Co^{3+}(aq) + 6NO_3^-(aq) \rightarrow Cd^{2+}(aq) + 2Co^{2+}(aq) + 6NO_3^-(aq)$

or

$2Co(NO_3)_3(aq) + Cd(s) \rightarrow 2Co(NO_3)_2(aq) + Cd(NO_3)_2(aq)$

b. *oxidation*: $SO_2(g) \rightarrow SO_4^{2-}(aq)$

$SO_2(g) + 2H_2O(\ell) \rightarrow SO_4^{2-}(aq)$

$SO_2(g) + 2H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 4H^+(aq)$

$SO_2(g) + 2H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

reduction: $Ag^+(aq) \rightarrow Ag(s)$

$Ag^+(aq) + e^- \rightarrow Ag(s)$

multiply oxidation rx. by1:

$SO_2(g) + 2H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

multiply reduction rx. by 2:

$2[Ag^+(aq) + e^- \rightarrow Ag(s)]$

$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

combine:

$SO_2(g) + 2H_2O(\ell) + 2Ag^+(aq) + \cancel{2e^-} \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + \cancel{2e^-} + 2Ag(s)$

balance and add spectator ion:

$SO_2(g) + 2H_2O(\ell) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2NO^{3-}(aq) + 2Ag(s)$

or

$2AgNO_3(aq) + SO_2(g) + 2H_2O(\ell) \rightarrow 2Ag(s) + H_2SO_4(aq) + 2HNO_3(aq)$

c. *oxidation*: $Al(s) \rightarrow Al(OH)_3(s)$

$Al(s) + 3H_2O(\ell) \rightarrow Al(OH)_3(s)$

$Al(s) + 3H_2O(\ell) \rightarrow Al(OH)_3(s) + 3H^+(aq)$

$Al(s) + 3H_2O(\ell) + 3OH^-(aq) \rightarrow Al(OH)_3(s) + 3H^+(aq) + 3OH^-(aq)$

$Al(s) + \cancel{3H_2O(\ell)} + 3OH^-(aq) \rightarrow Al(OH)_3(s) + \cancel{3H_2O(\ell)}$

$Al(s) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

$Al(s) + 3OH^-(aq) \rightarrow Al(OH)_3(s) + 3e^-$

reduction:

$CrO_4^{2-}(aq) \rightarrow Cr(OH)_3(s)$

$CrO_4^{2-}(aq) \rightarrow Cr(OH)_3(s) + H_2O(\ell)$

$CrO_4^{2-}(aq) + 5H^+(aq) \rightarrow Cr(OH)_3(s) + H_2O(\ell)$

$CrO_4^{2-}(aq) + 5H^+(aq) + 5OH^-(aq) \rightarrow Cr(OH)_3(s) + H_2O + 5OH^-(aq)$

$CrO_4^{2-}(aq) + 5H_2O(\ell) \rightarrow Cr(OH)_3(s) + H_2O(\ell) + 5OH^-(aq)$

$CrO_4^{2-}(aq) + 4H_2O(\ell) \rightarrow Cr(OH)_3(s) + 5OH^-(aq)$

$CrO_4^{2-}(aq) + 4H_2O(\ell) + 3e^- \rightarrow Cr(OH)_3(s) + 5OH^-(aq)$

multiply oxidation rx. by 1:

$Al(s) + 3OH^-(aq) \rightarrow Al(OH)_3(s) + 3e^-$

multiply reduction rx. by 1:

$CrO_4^{2-}(aq) + 4H_2O(\ell) + 3e^- \rightarrow Cr(OH)_3(s) + 5OH^-(aq)$

combine:

$Al(s) + 3OH^-(aq) + CrO_4^{2-}(aq) + 4H_2O(\ell) + \cancel{3e^-} \rightarrow Al(OH)_3(s) + 3e^- + Cr(OH)_3(s) + 5OH^-(aq)$

balance:

$Al(s) + CrO_4^{2-}(aq) + 4H_2O(\ell) \rightarrow Al(OH)_3(s) + Cr(OH)_3(s) + 2OH^-(aq)$

add spectator ions:

$Al(s) + 2Na^+(aq) + CrO_4^{2-}(aq) + 4H_2O(\ell) \rightarrow Al(OH)_3(s) + Cr(OH)_3(s) + 2Na^+(aq) + 2OH^-(aq)$

or

$Al(s) + Na_2CrO_4(aq) + 4H_2O(\ell) \rightarrow Al(OH)_3(s) + Cr(OH)_3(s) + 2NaOH(aq)$

37. a. $I_2(s) + 10HNO_3(aq) \rightarrow 2HIO_3(aq) + 10NO_2(g) + 4H_2O(\ell)$

b. The oxidation number for iodine is 0 in $I_2(s)$ and +5 in $HIO_3(aq)$. This is loss of 5 electrons or an oxidation. The oxidation number of nitrogen in $HNO_3(aq)$ is +5 and in $NO_2(g)$ is +4. This is gain of 1 electron or a reduction.

38. $C > A > D > B$

$B < D < A < C$

39. a. Calcium is a stronger reducing agent than iron. Therefore, the reaction where iron is the reducing agent is not a spontaneous reaction.

b. The reverse reaction will be spontaneous.

40. Oxidation numbers (left to right):

$Ca(s)$	$+ 2H_2O(\ell) \rightarrow$	$Ca(OH)_2(aq)$	$+ H_2(g)$
0	+1 −2	+2 −2 +1	0

Redox reaction: Ca is oxidized and H is reduced

41. a. This is a redox reaction, as the oxidation number of the oxidation number of silver changes from 0 to +1, an oxidation, and the oxidation number of oxygen changes from 0 to −2, a reduction.

b. $8CN^-(aq) + 4Ag(s) + O_2(g) + 2H_2O(\ell) \rightarrow 4Ag(CN)_2^-(aq) + 4OH^-(aq)$

42. a. *oxidation*: $I^-(aq) \rightarrow I_2(s)$

$2I^-(aq) \rightarrow I_2(s)$

$2I^-(aq) \rightarrow I_2(s) + 2e^-$

reduction: $Cr_2O_7^{2-} \rightarrow Cr^{3+}(aq)$

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}(aq)$

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$

$Cr_2O_7^{2-} + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$

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

multiply oxidation rx. by 3:

$3[2I^-(aq) \rightarrow I_2(s) + 2e^-]6I^-(aq) \rightarrow 3I_2(s) + 6e^-$

multiply reduction rx. by 1:

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

combine:

$6I^-(aq) + Cr_2O_7^{2-} + 14H^+(aq) + \cancel{6e^-} \rightarrow 3I_2(s) + \cancel{6e^-} + 2Cr^{3+}(aq) + 7H_2O(\ell)$

balance:

$6I^-(aq) + Cr_2O_7^{2-} + 14H^+(aq) \rightarrow 3I_2(s) + 2Cr^{3+}(aq) + 7H_2O(\ell)$

Ionic equations are usually done more readily by the half-reaction method.

b. *oxidation*: $Cl^-(aq) \rightarrow Cl_2(g)$

$2Cl^-(aq) \rightarrow Cl_2(g)$

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

reduction: $SeO_4^{2-}(aq) \rightarrow SeO_3^{2-}(aq)$

$SeO_4^{2-}(aq) \rightarrow SeO_3^{2-}(aq) + H_2O(\ell)$

$SeO_4^{2-}(aq) + 2H^+(aq) \rightarrow SeO_3^{2-}(aq) + H_2O(\ell)$

$SeO_4^{2-}(aq) + 2H^+(aq) + 2OH^-(aq) \rightarrow SeO_3^{2-}(aq) + H_2O(\ell) + 2OH^-(aq)$

$SeO_4^{2-}(aq) + 2H_2O(\ell) + 2e^- \rightarrow SeO_3^{2-}(aq) + H_2O(\ell) + 2OH^-(aq)$

multiply oxidation rx. by 1:

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

multiply reduction rx. by 1:

$SeO_4^{2-}(aq) + 2H_2O(\ell) + 2e^- \rightarrow SeO_3^{2-}(aq) + H_2O(\ell) + 2OH^-(aq)$

combine:

$2Cl^-(aq) + SeO_4^{2-}(aq) + 2H_2O(\ell) + \cancel{2e^-} \rightarrow Cl_2(g) + \cancel{2e^-} + SeO_3^{2-}(aq) + H_2O(\ell) + 2OH^-(aq)$

balance:

$2Cl^-(aq) + SeO_4^{2-}(aq) + H_2O(\ell) \rightarrow Cl_2(g) + SeO_3^{2-}(aq) + 2OH^-(aq)$

Ionic equations are usually done more readily by the half-reaction method.

c. *oxidation*: $SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq)$

$SO_3^{2-}(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq)$

$SO_3^{2-}(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 2H^+(aq)$

$SO_3^{2-}(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$

reduction: $MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$

$MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

$MnO_4^-(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

multiply oxidation rx. by 5:

$5[SO_3^{2-}(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-]$

$5SO_3^{2-}(aq) + 5H_2O(\ell) \rightarrow 5SO_4^{2-}(aq) + 10H^+(aq) + 10e^-$

multiply reduction rx. by 2:

$2[MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)]$

$2MnO_4^-(aq) + 16H^+(aq) + 10e^- \rightarrow 2Mn^{2+}(aq) + 8H_2O(\ell)$

combine:

$5SO_3^{2-}(aq) + 5H_2O(\ell) + 2MnO_4^-(aq) + 16H^+(aq) + \cancel{10e^-} \rightarrow 5SO_4^{2-}(aq) + 10H^+(aq) + \cancel{10e^-} + 2Mn^{2+}(aq) + 8H_2O(\ell)$

balance:

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(\ell)$

Ionic equations are usually done more readily by the half-reaction method.

d. *oxidation*: $Ag(s) \rightarrow Ag^+(aq)$

$Ag(s) \rightarrow Ag^+(aq) + e^-$

reduction: $NO_3^-(aq) \rightarrow NO(g)$

$NO_3^-(aq) \rightarrow NO(g) + 2H_2O(\ell)$

$NO_3^-(aq) + 4H^+(aq) \rightarrow NO(g) + 2H_2O(\ell)$

$NO_3^-(aq) + 4H^+(aq) + 4OH^-(aq) \rightarrow NO(g) + 2H_2O(\ell) + 4OH^-(aq)$

$NO_3^-(aq) + 4H_2O(\ell) \rightarrow NO(g) + 2H_2O(\ell) + 4OH^-(aq)$

$NO_3^-(aq) + 4H_2O(\ell) + 3e^- \rightarrow NO(g) + 2H_2O(\ell) + 4OH^-(aq)$

multiply oxidation rx. by 3:

$3[Ag(s) \rightarrow Ag^+(aq) + e^-]$

$3Ag(s) \rightarrow 3Ag^+(aq) + 6e^-$

multiply reduction rx. by 1:

$NO_3^-(aq) + 4H_2O(\ell) + 3e^- \rightarrow NO(g) + 2H_2O(\ell) + 4OH^-(aq)$

combine:

$3Ag(s) + NO_3^-(aq) + 4H_2O(\ell) \cancel{+ 3e^-} \rightarrow 3Ag^+(aq) \cancel{+ 3e^-} + NO(g) + 2H_2O(\ell) + 4OH^-(aq)$

balance:

$3Ag(s) + NO_3^-(aq) + 2H_2O(\ell) \rightarrow 3Ag^+(aq) + NO(g) + 4OH^-(aq)$

43.
- oxidation: loss of electrons; oxidation number increases
- oxidation and reduction: gain of electrons is numerically equal to the loss of electrons
- reduction: gain of electrons; oxidation number decreases

44. An increase in oxidation numbers indicates a loss of electrons, therefore an oxidation; a decrease in oxidation numbers indicates a gain of electrons, therefore a reduction. The change associated with the increase and the decrease can be balanced by finding the lowest common multiple of each change and multiplying the half reactions by the factor needed to obtain the common multiple.

45. The only difference occurs at the end of the reaction. All redox reactions are first balanced in acidic conditions, and, once balanced, the hydrogen ions are neutralized by adding the same number of hydroxide ions as there are hydrogen ions to both sides. The side with the hydrogen ions forms water molecules with the addition of the hydroxide ions, which can then reduce the water molecules in the overall reaction

46. Zn(s), Al(s) and Mg(s) could be used. All are stronger reducing agents than Cr(s). A reducing agent is oxidized or loses electrons. Each of these metals will lose electrons to the Cr^{3+} ions and these ions will be reduced to solid chromium.

47. a. In this general reaction A goes from an oxidation number of 0 to a positive oxidation number, and B goes from 0 to a negative oxidation number. Since there is an increase and a decrease in oxidation numbers, this is a redox reaction.

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

$Zn(s) + S(g) \rightarrow ZnS(s)$

b. A is the reducing agent, and B is the oxidizing agent.

48. Adding to both sides is required to maintain the balance of charge and mass that exists after the reaction was balanced in acidic conditions.

49. Answers should demonstrate students understand how oxidations and reductions have a major impact on the two examples cited from their research. Answers should include the redox reactions they are discussing and specifically their impact in industry, on health and safety and on the environment.

50. The electrons are written on the product side (right side) of the arrow in an oxidation half-reaction, because they are lost (produced) by the oxidation.

$FeSO_4(aq) + Zn(s) \rightarrow Fe(s) + ZnSO_4(aq)$

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

51. Student ideas are best: e.g.

oxidizing agent: Old AGE = Oxidizing Agent Gains Electrons

52. Since the reaction as written was not spontaneous, the reverse reaction would be spontaneous, as the reducing agent strengths would now be such that the reaction would occur spontaneously.

$2Ag(s) + Zn^{2+}(aq) \rightarrow 2Ag^+(aq) + Zn(s)$
not spontaneous

$2Ag^+(aq) + Zn(s) \rightarrow 2Ag(s) + Zn^{2+}(aq)$
spontaneous

Ag(s) is a weaker reducing agent than Zn(s) or alternatively $Ag^+(aq)$ is a stronger oxidizing agent than $Zn^{2+}(aq)$

53. The high electronegativity of the element fluorine causes the element to pull an electron from any atom in another molecule, causing it to take on a charge of -1, which is also the oxidation number of the element.

54. Step 1: The hydrogen atoms have an oxidation number of +1, so assign this value to each of the hydrogen atoms.

Step 2: The oxygen atoms have an oxidation number of −2, so assign this value to each of the oxygen atoms.

Step 3: Since the overall molecule is neutral, the sum of the oxidation numbers must be zero, so using x as the unknown, we solve $2(+1) + x + 3(-2) = 0$ to find the oxidation number of the sulfur.

55. Since aluminum is a stronger reducing agent than iron is, this would not be a good idea because the aluminum will be oxidized from solid Al to $Al^{3+}(aq)$, thus it would appear as if the solution were eating through the container

56. Answers should demonstrate students understand the process of various types of iron and steel and how redox plays a key role in these processes.

57. This report should be organized to describe the construction of the disc, materials used, electrochemistry that occurs. Rubric should include value for organization, completeness, as well as writing components including spelling and grammar.

58. Refer to the description of graphic organizers in Appendix A, pages 712-714. Key terms and concepts are summarized on page 623.

59. a. Na +1, Al +3, O −2, H +1, C +4

b. $NaAl(OH)_2CO_3(s) + 4HCl(aq) \rightarrow NaCl(aq) + AlCl_3(aq) + 3H_2O(\ell) + CO_2(g)$

c. This is not a redox reaction so the oxidation numbers are not of use here.

d. The reaction is an acid-base neutralization reaction.

60. a. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

$S(s) + O_2(g) \rightarrow SO_2(g)$

b. $KClO_3(s)$ is both the oxidizing agent and the reducing agent in the first reaction. S(s) is the reducing agent and $O_2(g)$ is the oxidizing agent in the second reaction

c. No element in potassium chlorate undergoes disproportionation. Each element appears only once in the formulas of the reactants, so each element undergoes reduction or oxidation, not both.

d. The safety match was invented in 1844 by Swedish chemist Gustaf Erik Pasch. Before then, matches were made using toxic yellow phosphorus, which also could light when struck on any surface. Pasch used less toxic and less flammable red phosphorus, which he removed from the match head and mixed with sandpaper on a special striking surface.

61. a. The first two steps are redox reactions, as the oxidation number of the sulfur atom changes from 0 to +4 in step 1 and from +4 to +6 in step 2. The oxygen atom changes from 0 to −2 in both steps 1 and 2. The third step is not a redox reaction as all oxidation numbers remain unchanged for all atoms in the molecules.

b. In step both of the first two steps, sulfur atoms are oxidized and oxygen atoms are reduced.

62. This is not a redox reaction. Add hydrogen ions to balance charges and then balance by inspection.

$Fe_2O_3(s) + 6H_2C_2O_4(aq) \rightarrow 2Fe(C_2O_4)_3{}^{3-}(aq) + 3H_2O(\ell) + 6H^+(aq)$

63. $Fe^{3+}(aq)$ is shown changing to $Fe^{2+}(aq)$ This is an oxidation half reaction. $Sn^{4+}(aq)$ is shown changing to $Sn^{2+}(aq)$. This is also an oxidation half reaction. Two oxidation half reactions cannot occur to one another.

64. a. $2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(\ell) + Al_2O_3(s)$

b. This reaction is a redox reaction. The oxidizing agent is Fe_2O_3 in which Fe changes in oxidation number from +3 to 0 in $Fe(\ell)$. The reducing agent is Al(s) which undergoes a change in oxidation number form 0 to +3 in $Al_2O_3(s)$.

65. Students should describe a procedure similar to Investigation 9.A, in which they would use single displacement reactions to decide whether tin or nickel is the better reducing agent. They would need samples of each metal, and solutions containing ions of the metals. They could then compare observations of tin metal in a solution of nickel ions, and nickel metal in a solution of tin ions. The metal that is the better reducing agent will react in a solution containing ions of the other metal.

66. $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

In this reaction, the carbon atom in carbon monoxide has an oxidation number of +2, and then goes to +4 in carbon dioxide. As a result, the increase in oxidation number indicates that the carbon is oxidized. The oxygen atom in molecular oxygen on the reactant side of the equation has an oxidation number of 0 and goes to −2 in carbon dioxide. As a result, the decrease in oxidation number indicates that the oxygen is reduced.

67. $Cr_2O_7{}^{2-}(aq) + C_2H_5OH(aq) \rightarrow Cr^{3+}(aq) + C_2H_4O_2(aq)$

Assign oxidation numbers:

$$\underset{+6\ \ -2}{Cr_2O_7^{2-}} + \underset{-2\ \ +1\ \ -2\ \ +1}{C_2H_5OH} \rightarrow \underset{+3}{Cr^{3+}} + \underset{0\ \ +1\ \ -2}{C_2H_4O_2}$$

Track electrons:

2(+6e⁻) = +12e⁻

$$Cr_2O_7^{2-} + C_2H_5OH \rightarrow Cr^{3+} + C_2H_4O_2$$

3(-4e⁻) = -12e⁻

Place coefficients determined by the gain and loss of electrons.

$2Cr_2O_7^{2-}(aq) + 3C_2H_5OH(aq) \rightarrow$
$Cr^{3+}(aq) + C_2H_4O_2(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the carbon and oxygen atoms remain as above:

$2Cr_2O_7^{2-}(aq) + 3C_2H_5OH(aq) \rightarrow$
$4Cr^{3+}(aq) + 3C_2H_4O_2(aq)$

Add water to balance the oxygen atoms.

$2Cr_2O_7^{2-}(aq) + 3C_2H_5OH(aq) \rightarrow$
$4Cr^{3+}(aq) + 3C_2H_4O_2(aq) + 11H_2O(\ell)$

Add hydrogen ions to balance the hydrogen atoms.

$16H^{+}(aq) + 2Cr_2O_7^{2-}(aq) + 3C_2H_5OH(aq) \rightarrow$
$4Cr^{3+}(aq) + 3C_2H_4O_2(aq) + 11H_2O(\ell)$

68. Assign oxidation numbers:

$$\underset{0}{Al}(s) + \underset{-3\ +1}{NH_4}\underset{+7\ -2}{ClO_4}(g) \longrightarrow \underset{+3\ \ -2}{Al_2O_3}(g) + \underset{+1\ -1}{HCl}(g) + \underset{0}{N_2}(g) + \underset{+1\ \ -2}{H_2O}(g)$$

Track electrons:

$$\overset{-3e^-}{Al(s)} + \overset{-3e^-}{NH_4}\underset{+8e^-}{ClO_4}(g) \longrightarrow Al_2O_3(g) + HCl(g) + N_2(g) + H_2O(g)$$

A total of 6 electrons were lost from the reactants and 8 electrons were gained by a reactant. The lowest common multiple of 6 and 8 is 24. To balance electrons, for the equations, a total of 24 electrons must be lost by the reactants and a total of 24 electrons must be gained by the products. Since NH_4ClO_4 contains the only element that is losing electrons, its coefficient must be 3. Since the nitrogen atom in NH_4ClO_4 is now losing a total of 9 electrons, the other 15 electrons must be lost from aluminum. Therefore, put a coefficient of 5 in front of the Al(s). This ratio of coefficients must remain the same while balancing.

$5Al(s) + 3NH_4ClO_4(g) \rightarrow$
$Al_2O_3(g) + HCl(g) + N_2(g) + H_2O(g)$

At a glance, it is obvious that the aluminum atoms in Al_2O_3 cannot be balanced with an integer because of the subscript 2. Therefore, double the original coefficients, and then balance. Balance in the order, Al, Cl, N, O, and H.

$10Al(s) + 6NH_4ClO_4(g) \rightarrow$
$5Al_2O_3(g) + 6HCl(g) + 3N_2(g) + 9H_2O(g)$

69. a. All are redox reactions.

b. Step 1 oxidizing agent $O_2(g)$; reducing agent $NH_3(g)$
Step 2 oxidizing agent $O_2(g)$; reducing agent NO(g)
Step 3 oxidizing agent and reducing agent is $NO_2(g)$

70. $2Al(s) + 3Hg^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Hg(\ell)$

$2Al(s) + 3Sn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Sn(s)$

Answers to Chapter 9 Self-Assessment Questions

(Student textbook pages 630-1)

1. d

2. a

3. d

4. d

5. a

6. a

7. d

8. e

9. d

10. e

11. a. N +5
b. C −3
c. Cl +7
d. Cr +6

12. Step 1. Identify oxygen in its form with a −2 oxidation number.

Step 2. Multiply this oxidation number by the number of oxygen atoms in the molecule: $-2 \times 8 = -16$.

Step 3. Since the molecule is neutral, the value calculated in step 2 must be equal to the total in positive oxygen numbers, all of which is attributed to the sulfur atoms: +16.

Step 4. Since 3 atoms of sulfur contribute to the total positive in oxygen numbers, take the total of +16 and divide by 3 to obtain the oxidation number of the sulfur atoms: $+\frac{16}{3}$

13. Carbon changes from −2 to +4.

14. *Oxidation*:

$Zn(s) + 2OH^{-}(aq) \rightarrow ZnO(s) + H_2O(\ell) + 2e^{-}$

Reduction:

$HgO(s) + H_2O(\ell) + 2e^{-} \rightarrow Hg(\ell) + 2OH^{-}(aq)$

Oxidizing agent: HgO(s);reducing agent: Zn(s)

15. a and c do not proceed spontaneously

b. *Balanced equation*:
$Sn(s) + CuSO_4(aq) \rightarrow SnSO_4(aq) + Cu(s)$
Ionic equation: $Sn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow$
$Sn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$
Net ionic equation:
$Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$

16. a. *oxidation*: $Zn(s) \rightarrow Zn^{2+}(aq)$

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

reduction: $NO_3^-(aq) \rightarrow NH_4^+(aq)$

$NO_3^-(aq) \rightarrow NH_4^+(aq) + 3H_2O(\ell)$

$NO_3^-(aq) + 10H^+(aq) \rightarrow NH_4^+(aq) + 3H_2O(\ell)$

$NO_3^-(aq) + 10H^+(aq) + 8e^- \rightarrow NH_4^+(aq) + 3H_2O(\ell)$

multiply oxidation rx. by 4:

$4[Zn(s) \rightarrow Zn^+(aq) + 2e^-]$

$4Zn(s) \rightarrow 4Zn^+(aq) + 8e^-$

multiply reduction rx. by 1:

$NO_3^-(aq) + 10H^+(aq) + 8e^- \rightarrow NH_4^+(aq) + 3H_2O(\ell)$

combine:

$4Zn(s) + NO_3^-(aq) + 10H^+(aq) \cancel{+ 8e^-} \rightarrow 4Zn^+(aq) \cancel{+ 8e^-} + NH_4^+(aq) + 3H_2O(\ell)$

balance:

$4Zn(s) + NO_3^-(aq) + 10H^+(aq) \rightarrow 4Zn^{2+}(aq) + NH_4^+(aq) + 3H_2O(\ell)$

b. *oxidation*:

$CN^-(aq) \rightarrow CNO^-(aq)$

$CN^-(aq) + H_2O(\ell) \rightarrow CNO^-(aq)$

$CN^-(aq) + H_2O(\ell) \rightarrow CNO^-(aq) + 2H^+(aq)$

$CN^-(aq) + H_2O(\ell) + 2OH^-(aq) \rightarrow CNO^-(aq) + 2H^+(aq) + 2OH^-(aq)$

$CN^-(aq) \cancel{+ H_2O(\ell)} + 2OH^-(aq) \rightarrow CNO^-(aq) + \cancel{2}H_2O(\ell)$

$CN^-(aq) + 2OH^-(aq) \rightarrow CNO^-(aq) + H_2O(\ell) + 2e^-$

reduction:

$CrO_4^{2-}(aq) \rightarrow Cr(OH)_3(s)$

$CrO_4^{2-}(aq) \rightarrow Cr(OH)_3(s) + H_2O(\ell)$

$CrO_4^{2-}(aq) + 5H^+(aq) \rightarrow Cr(OH)_3(s) + H_2O(\ell)$

$CrO_4^{2-}(aq) + 5H^+(aq) + 5OH^-(aq) \rightarrow Cr(OH)_3(s) + H_2O(\ell) + 5OH^-(aq)$

$CrO_4^{2-}(aq) + 5H_2O(\ell) \rightarrow Cr(OH)_3(s) + H_2O(\ell) + 5OH^-(aq)$

$CrO_4^{2-}(aq) + 4H_2O(\ell) + 3e^- \rightarrow Cr(OH)_3(s) + 5OH^-(aq)$

multiply oxidation rx. by 3:

$3[CN^-(aq) + 2OH^-(aq) \rightarrow CNO^-(aq) + H_2O(\ell) + 2e^-]$

$3CN^-(aq) + 6OH^-(aq) \rightarrow 3CNO^-(aq) + 3H_2O(\ell) + 6e^-$

multiply reduction rx. by 2:

$2[CrO_4^{2-}(aq) + 4H_2O(\ell) + 3e^- \rightarrow Cr(OH)_3(s) + 5OH^-(aq)]$

$2CrO_4^{2-}(aq) + 8H_2O(\ell) + 6e^- \rightarrow 2Cr(OH)_3(s) + 10OH^-(aq)$

combine:

$3CN^-(aq) + 6OH^-(aq) + 2CrO_4^{2-}(aq) + 8H_2O(\ell) \cancel{+ 6e^-} \rightarrow 3CNO^-(aq) + 3H_2O(\ell) \cancel{+ 6e^-} + 2Cr(OH)_3(s) + 10OH^-(aq)$

balance:

$3CN^-(aq) + 2CrO_4^{2-}(aq) + 5H_2O(\ell) \rightarrow 3CNO^-(aq) + 2Cr(OH)_3(s) + 4OH^-(aq)$

c. *oxidation*: $I^-(aq) \rightarrow I_2(s)$

$2I^-(aq) \rightarrow I_2(s)$

$2I^-(aq) \rightarrow I_2(s) + 2e^-$

reduction: $BrO_3^-(aq) \rightarrow Br^-(aq)$

$BrO_3^-(aq) \rightarrow Br^-(aq) + 3H_2O(\ell)$

$BrO_3^-(aq) + 6H^+(aq) \rightarrow Br^-(aq) + 3H_2O(\ell)$

$BrO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow Br^-(aq) + 3H_2O(\ell)$

multiply oxidation rx. by 3:

$3[2I^-(aq) \rightarrow I_2(s) + 2e^-]$

$6I^-(aq) \rightarrow 3I_2(s) + 6e^-$

multiply reduction rx. by 1:

$BrO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow Br^-(aq) + 3H_2O(\ell)$

combine:

$6I^-(aq) + BrO_3^-(aq) + 6H^+(aq) \cancel{+ 6e^-} \rightarrow 3I_2(s) \cancel{+ 6e^-} + Br^-(aq) + 3H_2O(\ell)$

balance:

$6I^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3I_2(s) + Br^-(aq) + 3H_2O(\ell)$

17. Since the sulfide ions have already gained two electrons to become the ion, they would not be able to gain more electrons, meaning that sulfide can no longer be reduced. Therefore, if it cannot go through a reduction, it cannot be an oxidizing agent.

18. a. Assign oxidation numbers:

$$\underset{+7\ -2}{MnO_4^-}(aq) + \underset{+2}{Fe^{2+}}(aq) \rightarrow \underset{+2}{Mn^{2+}}(aq) + \underset{+3}{Fe^{3+}}(aq)$$

Track electrons:

$$\overbrace{MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}}^{1(+5e^-) = +5e^-}(aq) + Fe^{3+}(aq)$$

$$MnO_4^-(aq) + \underbrace{Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}}_{5(-1e^-) = -5e^-}(aq)$$

Place coefficients determined by the gain and loss of electrons.

$MnO_4^-(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the carbon and oxygen atoms remain as above:

Already balanced.

Add water to balance the oxygen atoms.

$MnO_4^-(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 2H_2O(\ell)$

Add hydrogen ions to balance the hydrogen atoms.

$MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(\ell)$

b. Assign oxidation numbers:

$$\underset{+3}{Cr^{3+}}(aq) + \underset{+5\ -2}{ClO_3^-}(aq) \rightarrow \underset{+4\ -2}{ClO_2}(aq) + \underset{+6\ -2}{Cr_2O_7^{2-}}(aq)$$

Track electrons:

$$\overbrace{Cr^{3+}(aq) + ClO_3^-(aq) \rightarrow ClO_2(aq) + Cr_2O_7^{2-}(aq)}^{1(-3e^-) = -3e^-}$$
$$3(+1e^-) = +3e^-$$

Place coefficients determined by the gain and loss of electrons:

$Cr^{3+}(aq) + 3ClO_3^-(aq) \rightarrow ClO_2(aq) + Cr_2O_7^{2-}(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the carbon and oxygen atoms remain as above:

$2Cr^{3+}(aq) + 6ClO_3^-(aq) \rightarrow 6ClO_2(aq) + Cr_2O_7^{2-}(aq)$

Add water to balance the oxygen atoms:

$2Cr^{3+}(aq) + 6ClO_3^-(aq) + H_2O(\ell) \rightarrow 6ClO_2(aq) + Cr_2O_7^{2-}(aq)$

Add hydrogen ions to balance the hydrogen atoms:

$2Cr^{3+}(aq) + 6ClO_3^-(aq) + H_2O(\ell) \rightarrow 6ClO_2(aq) + Cr_2O_7^{2-}(aq) + 2H^+(aq)$

c. Assign oxidation numbers:

$$\underset{+5\ -2}{NO_3^-}(aq) + \underset{0}{Bi}(s) \rightarrow \underset{+3}{Bi^{3+}}(aq) + \underset{+4\ -2}{NO_2}(aq)$$

Track electrons:

$$\overbrace{NO_3^-(aq) + Bi(s) \rightarrow Bi^{3+}(aq) + NO_2(aq)}^{3(+1e^-) = +3e^-}$$
$$1(-3e^-) = -3e^-$$

Place coefficients determined by the gain and loss of electrons:

$3NO_3^-(aq) + Bi(s) \rightarrow Bi^{3+}(aq) + NO_2(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the carbon and oxygen atoms remain as above:

$3NO_3^-(aq) + Bi(s) \rightarrow Bi^{3+}(aq) + 3NO_2(aq)$

Add water to balance the oxygen atoms:

$3NO_3^-(aq) + Bi(s) \rightarrow Bi^{3+}(aq) + 3NO_2(aq) + 3H_2O(\ell)$

Add hydrogen ions to balance the hydrogen atoms:

$3NO_3^-(aq) + Bi(s) + 6H^+(aq) \rightarrow Bi^{3+}(aq) + 3NO_2(aq) + 3H_2O(\ell)$

To both sides, add the number of hydroxide ions equal to the number hydrogen ions.

$3NO_3^-(aq) + Bi(s) + 6H^+(aq) + 6OH^-(aq) \rightarrow Bi^{3+}(aq) + 3NO_2(aq) + 3H_2O(\ell) + 6OH^-(aq)$

Combine H^+ and OH^- to make water, where possible.

$3NO_3^-(aq) + Bi(s) + 6H_2O(\ell) \rightarrow Bi^{3+}(aq) + 3NO_2(aq) + 3H_2O(\ell) + 6OH^-(aq)$

Cancel water molecules that are on both sides of the equation and check to ensure that atoms of all elements are balanced.

$3NO_3^-(aq) + Bi(s) + 3H_2O(\ell) \rightarrow Bi^{3+}(aq) + 3NO_2(aq) + 6OH^-(aq)$

19. a. Assign oxidation numbers:

$$\underset{-2\ +1\ -2\ +1}{CH_3OH}(\ell) + \underset{0}{O_2}(g) \rightarrow \underset{0\ +1\ -2}{CH_2O}(\ell) + \underset{+1\ -2}{H_2O}(\ell)$$

Track electrons:

$$\overbrace{CH_3OH(\ell) + O_2(g) \rightarrow CH_2O(\ell) + H_2O(\ell)}^{2(-2e^-) = -4e^-}$$
$$1(+4e^-) = +4e^-$$

Place coefficients determined by the gain and loss of electrons:

$2CH_3OH(\ell) + O_2(g) \rightarrow 2CH_2O(\ell) + H_2O(\ell)$

Balance while ensuring that the ratios of the carbon and oxygen atoms remain as above:

$2CH_3OH(\ell) + O_2(g) \rightarrow 2CH_2O(\ell) + 2H_2O(\ell)$

b. Reducing agent: $CH_3OH(\ell)$; oxidizing agent: $O_2(g)$

20. Sample answers:

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

$Zn(s) + 2Au^+(aq) \rightarrow Zn^{2+}(aq) + 2Au(s)$

$Zn(s) + Ni^{2+}(aq) \rightarrow Zn^{2+}(aq) + Ni(s)$

21. In the smelting process, pig iron is extracted from an iron ore using CO(g) as a reducing agent. Redox reactions that occur include:

$2C(s) + O_2(g) \rightarrow 2CO(g)$

$Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$

$Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$

$FeO(s) + CO(g) \rightarrow Fe(\ell) + CO_2(g)$

22. The oxidation number of bromine decreases from 0 in $Br_2(\ell)$ to -1 in $2Br^-$ (aq). This is reduction. $Br_2(\ell)$ is an oxidizing agent. The oxidation number of chlorine increases from +3 in $ClO_2^-(aq)$ to +4 in $ClO_2(aq)$. This is oxidation. $ClO_2^-(aq)$ is a reducing agent.

23. $NH_4^+(aq) + 2O_2(g) \rightarrow NO_3^-(aq) + H_2O(\ell) + 2H^+(aq)$

24. To have an oxidation number change from +4 to +6, an increase of 2 means that the atom must have lost 2 electrons and therefore went through oxidation in the reaction. The substance is therefore a reducing agent.

25. $4Al(s) + 3NH_4ClO_4(s) \rightarrow 4Al_2O_3(s) + 3NH_4Cl(s)$

Al changes in oxidation number from 0 in Al(s) to +3 in Al_2O_3(s). This is an oxidation. Al(s) is the reducing agent. Cl changes in oxidation number from +7 in NH_4ClO_4(s) to −1 in NH_4Cl(s). This is a reduction. NH_4ClO_4(s) is the oxidizing agent

Chapter 10 Electrochemical Cells

Answers to Learning Check Questions

(Student textbook page 641)

1. The zinc atoms can reduce the copper(II) ions even though the two metals are not in contact because the electrons from the zinc travel to the electrodes via the connecting wire before coming into contact with the copper(II) ions in solution. Ions migrate through a salt bridge to maintain a balance of ionic charge.

2. The salt bridge in a galvanic cell allows movement of the anions and cations, as oxidation and reduction half-reactions occur. This prevents a build-up of charge in each half-cell.

3. At the anode, atoms are converted into cations by an oxidation reaction, and the electrode loses mass. At the cathode, cations are converted into atoms by a reduction reaction, and the electrode gains mass.

4. a. In a test tube, the zinc would be the reducing agent and the copper(II) ions would be the oxidizing agent.

   b. In a Daniell cell, the zinc would be the reducing agent and the copper(II) ions would be the oxidizing agent. In this cell, the two half-cell reactions are in separate compartments but the overall reaction is the same.

5. a. *Oxidation (anode):* $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

   Reduction (cathode): $Tl^+(aq) + 1e- \rightarrow Tl(s)$

   Overall cell reaction:

   $Sn(s) + 2Tl^+(aq) \rightarrow Sn^{2+}(aq) + 2Tl(s)$

   b. *Oxidation (anode):* $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$

   Reduction (cathode): $2H^+(aq) + 2e^- \rightarrow H_2(g)$

   Overall cell reaction:

   $Cd(s) + 2H^+(aq) \rightarrow Cd^{2+}(aq) + H_2(g)$

   Platinum is an inert electrode.

6. There is no double vertical line in the notation because there is no salt bridge in the lemon to separate the two half-cell reactions

(Student textbook 655)

7. The dry cell and Daniell cell are both galvanic cells. They are similar in that both use a zinc anode and the same reaction will occur at the anode. They differ as follows: The Daniell cell uses a copper cathode in a solution containing Cu^{2+} that is separated from the anode by a salt bridge. The dry cell uses a paste of MnO_2, $ZnCl_2$, NH_4Cl, and carbon black as the electrolyte between the two half-cells. The cathode in the dry cell is made of graphite. The reduction reactions at the cathodes in the two cells are different.

   Daniell cell: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

   Dry cell:
   $2MnO_2(s) + H_2O(\ell) + 2e^- \rightarrow Mn_2O_3(s) + 2OH^-(aq)$

8. A battery consists of several cells connected in series.

9. Different voltages are produced, because the reduction reactions are different and therefore the overall reactions and voltages will not be the same.

10. $2H^+(aq) + 2e^- \rightarrow H_2 \qquad E°_{anode} = 0.00\ V$

    $O_2(g) + 4H^+ + e^- \rightarrow 2H_2O(\ell) \qquad E°_{cathode} = +1.23\ V$

    $$E°_{cell} = E°_{cathode} - E°_{anode}$$
    $$= +1.23\ V - 0.00\ V$$
    $$= +1.23\ V$$

11. They are similar in that the overall reaction in a fuel cell is a combustion reaction. They are different because fuel cell reactions convert the energy in the fuel directly into electrical energy. A combustion reaction converts the energy in the fuel into heat, which is used to drive an engine or generator to produce electricity. Fuel cell reactions are more efficient than combustion reactions and they are cleaner. The major "waste" product produced by fuel cells is water.

12. Electrical energy can split water into hydrogen and oxygen. Hydrogen can also be chemically removed from hydrocarbons by reforming. Additionally, wind and solar energies are used to produce the hydrogen. Research is also under way to develop fuel cells that have internal reformers.

(Student textbook page 658)

13. $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $E°anode = -0.45$ V

$O_2(g) + 2H_2O(\ell) + 4e^- \rightarrow 4OH^-(aq)$

$E°_{cathode} = +0.40$ V

$$\begin{aligned} E°_{cell} &= E°_{cathode} - E°_{anode} \\ &= +0.40 \text{ V} - (-0.45 \text{ V}) \\ &= +0.85 \text{ V} \end{aligned}$$

14. The standard cell potentials are based on 1.0 mol/L concentrations and SATP conditions. The concentrations and conditions in the environment where corrosion occurs will be different. The cell potentials will be different to those calculated.

15. Aluminum is a stronger reducing agent than iron is and will act as the anode, protecting the iron cathode.

16. The proximity to the Atlantic ocean results in a climate where salt-laden fog and generally humid conditions are common. This provides an electrolyte to connect dissimilar metals, making corrosion more common

17. The tin covers the steel so that water and oxygen from the environment cannot reach the steel. This keeps the iron in the steel from corroding.

18. If the can is scratched, water and oxygen from the environment can now reach the steel under the tin. The iron in the steel now corrodes faster in contact with the tin than it would on its own. This happens because tin is a weaker reducing agent and less reactive than iron and acts as a cathode in each miniature galvanic cell on the surface of the can. The tin provides a large area of available cathodes for the small galvanic cells, and iron acts as the anode of each cell, and thus rusting is facilitated.

(Student textbook page 664)

19. A galvanic cell converts chemical energy into electrical energy, whereas an electrolytic cell converts electrical energy into chemical energy.

20. a. Reaction at anode: $2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

b. Reaction at cathode: $Ca^{2+}(\ell) + 2e^- \rightarrow Ca(\ell)$

c. $CaCl_2(\ell) \rightarrow Ca(\ell) + Cl_2(g)$

21. a. Reaction at negative electrode:
$Li^+(\ell) + 1e^- \rightarrow Li(\ell)$

b. Reaction at positive electrode:
$2Br^-(\ell) \rightarrow Br_2(\ell) + 2e^-$

c. $2Li^+(\ell) + 2Br^-(\ell) \rightarrow 2Li(\ell) + Br_2(\ell)$

22. The negative sign means that the reaction is not spontaneous. This negative value represents the minimum potential difference that you would have to apply from an external power source to drive the cell reaction for the electrolytic cell.

23. The current in the electrolytic cell flows in one direction. Because alternating current constantly changes direction, it would not be able to drive the redox reaction in the electrolytic cell.

24. Sodium is solid at standard temperatures and will neither conduct an electric current nor be purified by electrolysis in an aqueous solution. Thus, Davy heated the salts until they melted and applied electrolysis to the molten solids to isolate sodium.

Answers to Caption Questions

Figure 10.5 (Student textbook page 637): $Cu^{2+}(aq)$ ions gain electrons and deposit on the copper electrode. This causes the concentration of $Cu^{2+}(aq)$ to decrease. The Zn(s) atoms lose electrons, and $Zn^{2+}(aq)$ ions go into solution, causing the concentration of $Zn^{2+}(aq)$ to increase.

Figure 10.7 (Student textbook page 638): Solid platinum is a very weak reducing agent, so it is less likely to participate in the reaction.

Figure 10.25 (Student textbook page 662): The non-standard reduction potentials for the reactions that occur at each electrode are as follows:

Oxidation:
$2H_2O(\ell) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ $-E = -0.82$ V

Reduction:
$2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $E = -0.42$ V

The number of electrons gained during reduction must be the same as the number of electrons lost during oxidation. The equations show that the reduction reaction occurs twice for every oxidation reaction. Therefore, 2 mol $H_2(g)$ is produced for every 1 mol $O_2(g)$.

Answers to Section 10.1 Review Questions

(Student textbook page 648)

1. It is important to keep each half-cell separate so that an instantaneous reaction does not occur, "short-circuiting" the voltmeter.

2. The salt bridge is needed to complete the circuit so that the cell can operate. The electrolyte in the salt bridge will have negative ions flow towards the anode and positive ions flow towards the cathode. This movement of ions is responsible for completing the flow of electricity in the cell. The ends of the salt bridge are plugged with cotton or glass wool to avoid the mixing of materials from the half-cell and to stop the flow of electrolytes into the half-cells.

3. Graphic organizer must display a basic understanding of how to determine which half-reaction will be the reduction and which will be the oxidation, based on the location of each half-reaction in the standard reduction potentials chart. The answer must then also demonstrate an understanding that a spontaneous reaction in a galvanic cell will have a positive cell potential.

4. a. net ionic equation:

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

$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $\quad E°_{anode} = -0.26\ V$

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E°_{cathode} = +0.34\ V$

$$\begin{aligned} E°_{cell} &= E°_{cathode} - E°_{anode} \\ &= +0.34\ V - (-\ 0.26\ V) \\ &= +\ 0.60\ V \end{aligned}$$

b. $6H_2O(\ell) + 3O_2(g) + 12e^- \rightarrow 12(OH)^-(aq)$ $\quad E°_{anode} = +0.40\ V$

$4Au^{3+}(aq) + 12e^- \rightarrow 4Au(s)$ $\quad E°_{cathode} = +1.50\ V$

$$\begin{aligned} E°_{cell} &= E°_{cathode} - E°_{anode} \\ &= +1.50\ V - 0.40\ V \\ &= +\ 1.10\ V \end{aligned}$$

c. net ionic equation:

$3Cu(s) + 2Al^{3+}(aq) \rightarrow 3Cu^{2+}(aq) + 2Al(s)$

$3Cu^{2+}(aq) + 6e^- \rightarrow 3Cu(s)$ $\quad E°_{anode} = +0.34\ V$

$2Al^{3+}(aq) + 6e^- \rightarrow 2Al(s)$ $\quad E°_{cathode} = -1.66\ V$

$$\begin{aligned} E°_{cell} &= E°_{cathode} - E°_{anode} \\ &= -1.66\ V - 0.34\ V \\ &= -\ 2.00\ V \end{aligned}$$

5. a. and b. are spontaneous since these reactions have a positive cell potential

6. It would not be possible, as the strongest oxidizing agent is fluorine gas and the strongest reducing agent is lithium, which together would produce a cell potential of 5.91 V.

7. The cell potential describes the potential difference between two electrodes of a cell, or the amount of energy on a charge as it moves between two electrodes. The cell potential is dependent on both the anode and the cathode used. The standard reduction potential is a measure of the amount of energy for only the reduction half of cell. Since reduction cannot happen without oxidation, reduction potentials are measured against a standard reference, the hydrogen half-cell, which is set at a reduction potential of 0.00V.

8. +0.99 V

Oxidation (anode): $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Reduction (cathode): $Pd^{2+}(aq) + 2e^- \rightarrow Pd(s)$

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E°_{anode} = -\ 0.76\ V$

$Pd^{2+}(aq) + 2e^- \rightarrow Pd(s)$ $\quad E°_{cathode} = X$

$E°_{cell} = E°_{cathode} - E°_{anode}$

$+1.75\ V = X - (-\ 0.76\ V)$

$X = +\ 0.99\ V$

9. a. *Oxidation:* $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Reduction: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$

Overall cell reaction:

$Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$

b. +1.56 V

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E°_{anode} = -\ 0.76\ V$

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $\quad E°_{cathode} = +\ 0.80$

$$\begin{aligned} E°_{cell} &= E°_{cathode} - E°_{anode} \\ &= +\ 0.80\ V - (-0.76\ V) = +\ 1.56\ V \end{aligned}$$

c. Zn(s) is the reducing agent; $Ag^+(aq)$ is the oxidizing agent

d. The Ag cathode increases in mass; the Zn anode decreases in mass

10. The electrode that participates in the redox reaction will either gain mass (if it is the cathode) or lose mass (if it is the anode) as metallic atoms gain or lose electrons as the cell operates. An inert electrode does not participate in the reaction, thus will not change in mass. It simply gives the half-reaction a location to occur as electrons move in the cell. This type of electrode is used when either the oxidizing agent or the reducing agent is not a solid.

11. Each half-cell potential would have been assigned a different value, but when used in calculating the overall cell potential, there would have been no difference in the calculated value

12. Graphic organizers should include steps of how to find the oxidizing agent, reducing agent and how to calculate the cell potential from the two half-reaction potentials. Encourage students to be as creative as possible in answering this question.

Answers to Section 10.2 Review Questions
(Student textbook page 659)

1. The oxidizing agent is manganese dioxide. The reducing agent is zinc.

2. The top of a commercial 1.5 V dry cell is marked with a plus sign because the graphite cathode protrudes there.

3. **a.** *Oxidation:*
$Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(\ell) + 2e^-$
Reduction:
$Ag_2O(s) + H_2O(\ell) + 2e^- \rightarrow 2Ag(s) + 2OH^-(aq)$
Overall cell reaction:
$Zn(s) + Ag_2O(s) \rightarrow ZnO(s) + 2Ag(s)$

b. anode is $Zn(s)$; cathode is $Ag_2O(s)$

4. Mercury is produced which is poisonous and a hazard to environmental life.

5. 4 dry cells at 1.5 V each

6. The potential of a cell depends only on the substances used for the anode and the cathode, not on the amounts of the substances used. A cell with a gold anode and an iron cathode would produce a cell potential of 1.95 V. If three of these cells were combined in series, a potential of almost 6 V would be produced. If only small quantities of these substances were used, the battery would be small. A large cell that produces only 1.5 V of potential contains larger quantities of the reactants required for the redox reaction. A large cell would be able to provide an electrical current over a longer period of time.

7. The zinc container of the dry cell is involved with the oxidation half-reaction. As zinc is oxidized, it forms ions and dissolves into the electrolyte paste. Therefore, the container has less and less zinc as the reaction progresses.

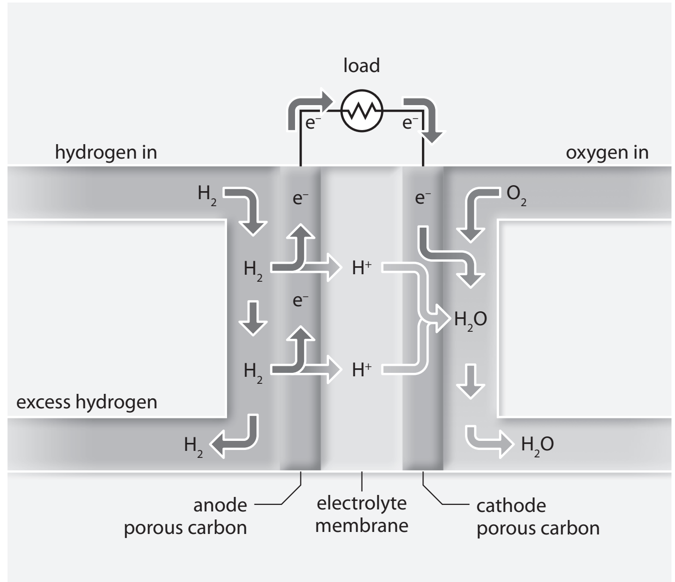
8. *Oxidation (anode):* $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
Reduction (cathode): $Ag^+(aq) + 1e^- \rightarrow Ag(s)$
Overall cell reaction:
$Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$
The design can be similar to a dry cell. The Fe is the anode and the Ag is the cathode.

9. Organizer should show that fuel cells and dry cells are alike in that they dispense with the need for an aqueous solution. In a dry cell, the electrolytes are in a paste, while fuel cells tend to rely on solid oxides, though the hydrogen fuel cell uses oxygen and hydrogen gases.

One of the chief differences of fuel cells and dry cells, however, is that the main waste of fuel cells can be recycled, such as water in the case of the hydrogen cell, whereas the waste of many depleted batteries involves toxic heavy metal cations.

10. The main benefit of fuel cells for space use is the possibility of recycling waste to re-start, reducing the need to re-supply with fuel or the store and remove fuel waste. Second, many fuel cells operate at higher efficiencies, allowing for more energy to be produced overall. Finally, the production of water helps to reduce risks of fuel combustion.

11. **a.**



b. *Oxidation (anode):* $H_2(g) \rightarrow 2H^+(s) + 2e^-$
Reduction (cathode):
$O_2(g) + 4H^+(s) + 4e^- \rightarrow 2H_2O(\ell)$

c. $H^+(aq)$ ions pass through the solid polymer electrolyte membrane but the negative electrons do not. The electrons move from the carbon anode, which is coated with platinum, through the external circuit to the cathode. Oxygen combines with the electrons and $H^+(aq)$ at the cathode to form $H_2O(\ell)$. The overall reaction is $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$

12. The use of salt lowers the melting point of snow, causing more water to be mobile and reactive. Furthermore, the salt also provides electrolytes required for the salt bridge.

13. Zinc is a more reactive metal than iron, as noted by its cell potential, so it will react before iron does.

14. **a.** Two metals that do not react in the presence of oxygen and water are silver and gold, as noted by their cell potentials being higher than that of the oxygen and water half-reaction.

b. Since these metals are quite resistant to corrosion they are useful for industrial processes that take place under extremely adverse conditions that promote corrosion, such as high temperatures or acidic environments, like the aerospace or petroleum industry. Also these metals are valued in jewellery since they are permanent.

15. a. The tarnish contains Ag_2O and/or Ag_2S. Treating with a stronger reducing agent than Ag will reverse the reaction.

b. $3Ag_2S(s) + 2Al(s) \rightarrow 6Ag(s) + Al_2S_3(s)$

Oxidation half-reaction: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

Reduction half-reaction: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$

Answers to Section 10.3 Review Questions

(Student textbook page 669)

1. a. The anode lead; the cathode is lead(IV) oxide, $PbO_2(s)$.

b. The sulfate ion, $SO_4{}^{2-}(aq)$ is a part of the electrolyte. It is required to combine with the $Pb^{2+}(aq)$ to precipitate $PbSO_4(s)$.

2. a. As a galvanic cell, the anode was negative and the cathode was positive. As an electrolytic cell, the anode is positive and the cathode is negative.

b. The flow of anions and cations will be reversed in the salt bridge.

3. a. The half-reactions are:

Reduction (occurs at the cathode):
$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$

Oxidation (occurs at the anode):
$2I^-(aq) \rightarrow I_2(s) + 2e^-$

$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s) \qquad E^\circ_{cathode} = -0.04\ V$

$I_2(s) + 2e^- \rightarrow 2I^-(aq) \qquad E^\circ_{anode} = +0.54\ V$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
$$= -0.04\ V - (0.54\ V) = -0.58\ V$$

The reaction has a negative standard cell potential; therefore, it is non-spontaneous under standard conditions.

b. The half-reactions are:

Reduction (cathode): $Ag^+(aq) + e^- \rightarrow Ag(s)$

Oxidation (anode):
$H_2SO_3(aq) + H_2O(\ell) \rightarrow SO_4{}^{2-}(aq) + 4H^+(aq) + 2e^-$

$SO_4{}^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O(\ell) \qquad E^\circ_{anode} = 0.17\ V$

$Ag^+(aq) + e^- \rightarrow Ag(s) \qquad E^\circ_{cathode} = +0.80\ V$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
$$= +0.80\ V - (0.17\ V)$$
$$= +0.63\ V$$

The reaction has a positive standard cell potential; therefore, it is spontaneous under standard conditions.

4. On discharging, the half-reactions are

$Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2(s) + 2e^-$

$NiO(OH)(s) + H_2O(\ell) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$

On recharging, therefore, the half-reactions are

$Cd(OH)_2(s) + 2e^- \rightarrow Cd(s) + 2OH^-(aq)$

$Ni(OH)_2(s) + OH^-(aq) \rightarrow NiO(OH)(s) + H_2O(\ell) + e^-$

To balance the charges, multiply the oxidation equation by 2:

$2Ni(OH)_2(s) + 2OH^-(aq) \rightarrow 2NiO(OH)(s) + 2H_2O(\ell) + 2e^-$

The overall reaction for recharging a nicad battery can then be written as

$Cd(OH)_2(s) + 2Ni(OH)_2(s) \rightarrow Cd(s) + 2NiO(OH)(s) + 2H_2O(\ell)$

5. > 12 V

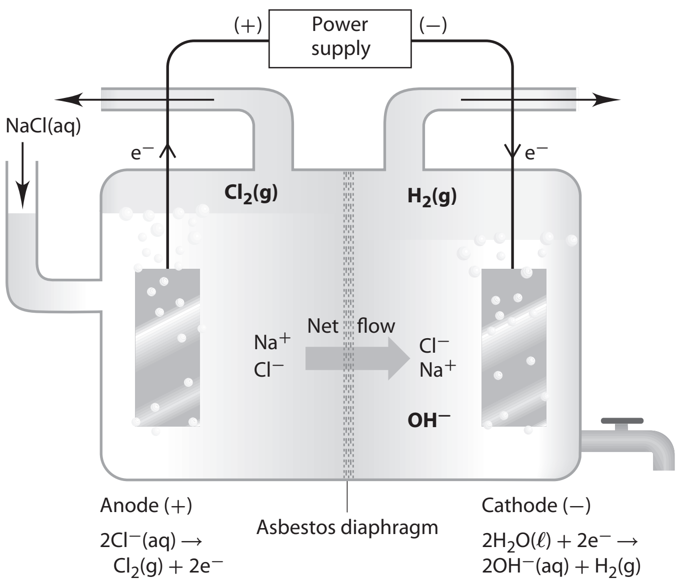
6. The reaction in an electrolytic cell is non-spontaneous; it requires an input of energy to occur. The energy is in the form of electrical energy. The external source of electrons forces electrons onto one electrode, which becomes negative relative to the other electrode. This drives the movement of ions and non-spontaneous redox reactions.

7. There are advantages and disadvantages to lead-acid batteries. Lead-acid batteries have been available for almost 150 years, so the technology is stable and has been improved on over the years. Modern batteries are completely sealed to minimize one of the potential risks, which is the possibility of a chemical burn if someone comes in contact with the acid in the battery. Although the charging process is never perfect, lead-acid batteries are long-lasting and if handled properly, much of the lead (up to 97%) can be recycled during disposal. On the negative side, lead is heavy, which decreases the energy efficiency of the car, and if disposed of improperly, lead will contaminate soil and water. If lead-acid batteries are cracked, the sulfuric acid inside will leak and may cause burns. In addition, if improperly charged or if an electrical short-circuit takes place, hydrogen gas that is produced when the battery is charged could be ignited if sparked, leading to a fire or an explosion.

8. a. The main products are sodium hydroxide and chlorine gas.

b. Sodium hydroxide is used in the pulp and paper industry to break down the lignin in wood and in the making of soap and detergents. Chlorine is used as a disinfectant, and to make laundry bleach, hydrochloric acid and to bleach pulp in paper making.

9.



The chlor-alkali cell is an electrolytic cell that oxidizes chloride ions, $Cl^-(aq)$ from brine (salt water) and reduces water, forming hydrogen gas and hydroxide ions.

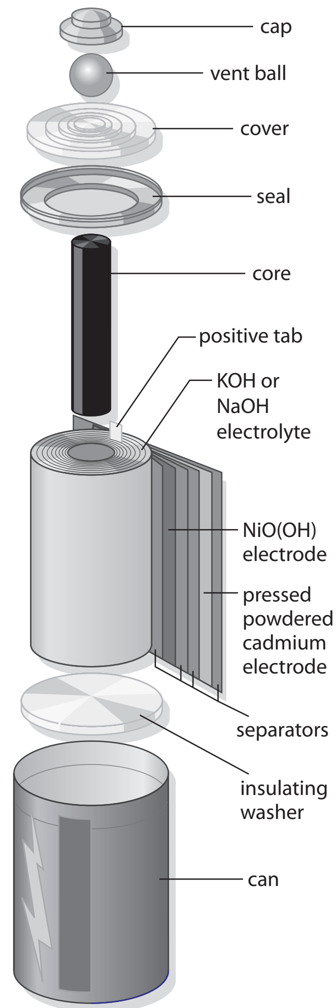
Oxidation: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

Reduction: $2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

The salt water enters the cell, and a battery provides the potential required to electrolyze the salt water. The products of the electrolysis are removed from the cell—the chlorine gas from the anode, the hydrogen gas from the cathode, and the aqueous sodium hydroxide is isolated from the electrolyte.

10. In theory this is correct. Practical considerations prevent this reversal in many cases. For example, when gases are involved, it is difficult to maintain contact with the electrodes. These are problems that chemical engineers try to solve.

11.



12. In the Down's cell, a direct current is passed through a mixture of molten $NaCl(\ell)$ and $CaCl_2(\ell)$.

$NaCl(\ell) \rightarrow Na^+(\ell) + Cl^-(\ell)$

$CaCl_2(\ell) \rightarrow Ca^{2+}(\ell) + 2Cl^-(\ell)$

$Na^+(\ell)$ is a stronger oxidizing agent than $Ca^{2+}(\ell)$. At the negative iron cathode, the reaction is $Na^+(\ell) + 1e^- \rightarrow Na(\ell)$

At the positive carbon anode, the reaction is $2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

Answers to Practice Problems

For full solutions to Practice Problems, see Part B of this Solutions Manual.

(Student textbook page 641)

1. Magnesium is the anode and zinc is the cathode.
2. Oxidation occurs at the magnesium electrode and reduction occurs at the zinc electrode.

3. The diagram should be similar to the one in the previous Sample Problem, but with magnesium instead of aluminum and zinc instead of nickel. Electron flow will be from the magnesium electrode to the zinc electrode.

4. Magnesium ions will flow into the magnesium half-cell, and zinc ions will flow onto the zinc electrode in the zinc half-cell.

5. $Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$

6. The anode is nickel and the cathode is silver.

7. Oxidation occurs at the nickel electrode and reduction occurs at the silver electrode.

8. The diagram should be similar to the one in the previous Sample Problem, but with nickel instead of aluminum (anode) and silver instead of nickel (cathode). Electron flow will be from the nickel electrode to the silver electrode.

9. $Ni^{2+}(aq)$ will flow into the nickel half-cell, and $Ag^{+}(aq)$ will flow onto the silver electrode in the silver half-cell.

10. $Ni(s) + 2Ag^{+}(aq) \rightarrow 2Ag(s) + Ni^{2+}(aq)$

(Student textbook page 647)

11. a. Oxidation: $2Br^{-}(aq) \rightarrow Br_2(\ell) + 2e^{-}$
Reduction: $Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(aq)$
b. $Br_2(\ell) + 2e^{-} \rightarrow 2Br^{-}(aq)$ $E°_{anode} = +1.07$ V
$Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(aq)$ $E°_{cathode} = +1.36$
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= 1.36 \text{ V} - (1.07 \text{ V}) = +0.29 \text{ V}$
The reaction is spontaneous in the direction written.

12. a. Oxidation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^{-}$
Reduction: $Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$
b. $Mg^{2+}(aq) + 2e^{-} \rightarrow Mg(s)$ $E°_{anode} = -2.37$ V
$Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$ $E°_{cathode} = +0.80$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= 0.80 \text{ V} - (-2.37 \text{ V}) = +3.17 \text{ V}$
The reaction is spontaneous in the direction written.

13. a. Oxidation: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^{-}$
Reduction: $2H^{+} + 2e^{-} \rightarrow H_2(g)$
b. $Sn^{2+}(aq) + 2e^{-} \rightarrow Sn(s)$ $E°_{anode} = -0.14$ V
$2H^{+}(aq) + 2e^{-} \rightarrow H_2(g)$ $E°_{cathode} = 0.00$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode} = 0.00 \text{ V} - (-0.14)$
$= 0.14 \text{ V}$
The reaction is spontaneous in the direction written.

14. a. Oxidation: $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^{-}$
Reduction: $Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$
b. $Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$ $E°_{anode} = -0.74$ V
$Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$ $E°_{cathode} = +0.80$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= 0.80 \text{ V} - (-0.74 \text{ V}) = +1.54 \text{ V}$
The reaction is spontaneous in the direction written.

15. a. *Oxidation:* $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$
Reduction: $Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$
b. $Fe^{2+}(aq) + 2e^{-} \rightarrow Fe(s)$ $E°_{anode} = -0.45$ V
$Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$ $E°_{cathode} = -0.74$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= -0.74 \text{ V} - (-0.45 \text{ V}) = -0.29 \text{ V}$
The reaction is not spontaneous in the direction written.

16. a. *Oxidation:* $Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$
Reduction: $Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$
b. $Al^{3+}(aq) + 3e^{-} \rightarrow Al(s)$ $E°_{anode} = -1.66$ V
$Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$ $E°_{cathode} = -0.76$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= -0.76 \text{ V} - (-1.66 \text{ V}) = +0.90 \text{ V}$
The reaction is spontaneous in the direction written.

17. a. *Oxidation:* $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$
Reduction: $Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$
b. $Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$ $E°_{anode} = -0.76$ V
$Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$ $E°_{cathode} = +0.80$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= 0.80 \text{ V} - (-0.76 \text{ V}) = +1.56 \text{ V}$
The reaction is spontaneous in the direction written.

18. a. *Oxidation:* $Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$
Reduction: $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$
b. $Al^{3+}(aq) + 3e^{-} \rightarrow Al(s)$ $E°_{anode} = -1.66$ V
$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$ $E°_{cathode} = +0.34$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= 0.34 \text{ V} - (-1.66 \text{ V}) = +2.00 \text{ V}$
The reaction is spontaneous in the direction written.

19. a. *Oxidation:* $Ag(s) \rightarrow Ag^{+}(aq) + 1e^{-}$
Reduction: $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$
b. $Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$ $E°_{anode} = +0.80$ V
$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$ $E°_{cathode} = +0.34$ V
c. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= 0.34 \text{ V} - (0.80 \text{ V}) = -0.46 \text{ V}$
The reaction is not spontaneous in the direction written.

20. a. *Oxidation:* $2I^-(aq) \rightarrow I_2(s) + 2e^-$

Reduction: $Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$

b. $I_2(s) + 2e^- \rightarrow 2I^-(aq) \qquad E°_{anode} = +0.54\ V$

$Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq) \qquad E°_{cathode} = +1.07\ V$

c. $E°_{cell} = E°_{cathode} - E°_{anode}$

$= 1.07\ V - (0.54\ V) = +0.53\ V$

The reaction is spontaneous in the direction written.

Answers to Chapter 10 Review Questions

(Student textbook pages 677-81)

1. d

2. a

3. c

4. b

5. b

6. d

7. e

8. b

9. d

10. c

11. a

12. c

13. a

14. a

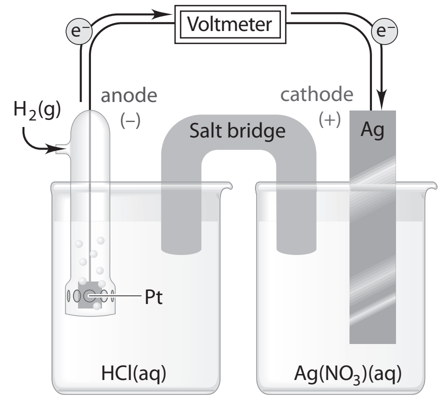
15. The quantity of electricity depends on the current generated. The current generated depends on the quantity and identity of substances involved in the electrochemical reaction. The batteries are different sizes and they may contain different materials. The larger battery is will last longer and produce a greater overall quantity of electricity.

16. The electrolyte is a paste that can be enclosed in a container. This is more convenient than the galvanic cell with its salt bridge and aqueous solutions.

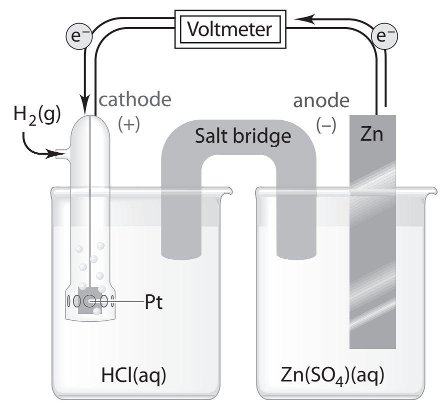
17. Using a sodium chloride solution in the salt bridge will eventually result in some chloride ions to migrating into the silver half-cell. This will cause a solid precipitate of silver chloride to form. As a result, the ionic concentration in the silver half-cell will decrease and thus the performance of the cell will be affected. To correct this, the electrolyte used in the salt bridge should be changed to an aqueous solution of sodium nitrate to avoid the formation of any precipitates as the electrolyte eventually migrates into the half-cells.

18. It is simply reversing the direction of the electron flow so that the reaction occurs in the opposite direction.; a reduction reaction becomes an oxidation reaction, and vice versa. The sign of the standard cell potential indicates this.

19. a.



b.



20. The standard half-cell potentials are measured using the hydrogen half-cell as a reference. In the equation $E°_{cell} = E°_{cathode} - E°_{anode}$, the hydrogen half-cell is given a value of zero. Any measured voltage would therefore belong to the non-hydrogen half of the cell.

21. Electrolyzing aqueous solutions often results in unwanted products due to redox reactions involving water.

22. Since zinc is a stronger reducing agent than is copper, the zinc will undergo oxidation when attached to a copper pipe, and thus can be used to protect a copper pipe.

23. Yes, as when written in the opposite direction, the cathode and anode interchange and the result will be a positive cell potential, which means the reaction is spontaneous.

24. a. Electrodes are the conductors that carry electrons into and out of the cell. Oxidation and reduction occur at the electrodes.

b. Electrolytes are substances that dissolve in water to form ions that can move and conduct electricity by the movement of ions.

c. The external voltage is a source of electricity that is included in the external circuit of an electrolytic cell.

25. $Mg(s) \mid Mg^{2+}(aq) \parallel Cd^{2+}(aq) \mid Cd(s)$

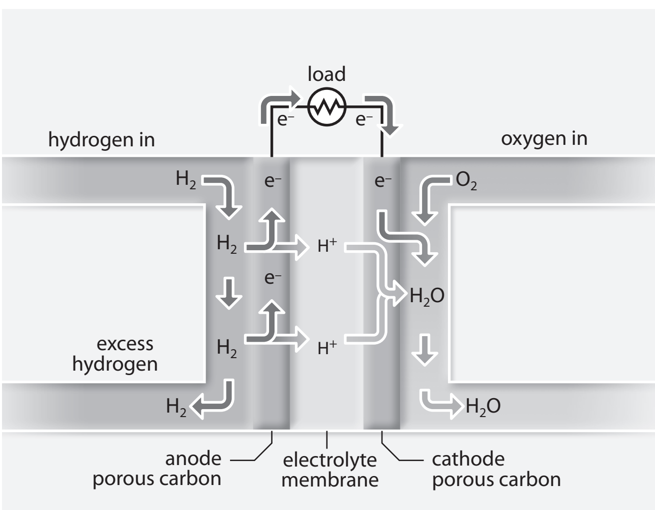
26. At standard temperatures, the NaCl(s) does not conduct electricity, therefore electrolysis cannot purify the metal. Instead, the salt can be heated until is becomes a molten solid. At this high temperature, the molten solid will conduct electricity and thus electrolysis can be used to purify the metal.

27.

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

Since copper is less reactive than iron, copper will act as the cathode and the iron will act as the anode of electrochemical cells at all points of contact between the two metals.

28.



c. $H^+(aq)$ ions pass through the solid polymer electrolyte membrane but the negative electrons do not. The electrons move from the carbon anode, which is coated with platinum, through the external circuit to the cathode. Oxygen combines with the electrons and $H^+(aq)$ at the cathode to form $H_2O(\ell)$.

The overall reaction is $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$.

29. $Br_2(\ell)$, $H^+(aq)$, $Co^{3+}(aq)$, $Zn^{2+}(aq)$

30. a. An actual galvanic cell could be set up using a Sn(s) anode in a $Sn(NO_3)_2(aq)$ solution, a Pb(s) cathode in a $Pb(NO_3)_2(aq)$ solution, connected externally with conducting wire with a voltmeter in parallel, and the cells connected with a salt bridge of $KNO_3(aq)$.

Alternatively, use the standard half-cell reduction potentials.

$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s) \qquad E°_{anode} = -0.14\ V$

$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s) \qquad E°_{cathode} = -0.13\ V$

b. $E°_{cell} = E°_{cathode} - E°_{anode}$

$= -0.13\ V - (-0.14\ V) = +0.01\ V$

c. This cell is unlikely to find any practical uses because the cell voltage is so low.

31. a. The oxidizing agent is $PbO_2(s)$ and the reducing agent is Pb(s).

b. During recharging, $PbSO_4(s)$ is both the oxidizing agent and the reducing agent.

32. It would not be possible, as the strongest oxidizing agent is fluorine gas and the strongest reducing agent is lithium, which together would produce a cell potential of 5.91 V.

33. As more and more ions migrate into solution from the anode, positive charge will build up in the aqueous solution. Eventually, due to like charge repulsion, the increased charge in the solution will start to hinder the movement of more positive ions into the solution. At the same time, the ions in the reduction half-cell are being removed from solution as they migrate onto the cathode. Over time, this will decrease the concentration of ions in this half-cell to a point where fewer and fewer ions are available to migrate to the electrode. Both of these events occurring simultaneously in the two half-cells will cause a drop in the voltage of the cell.

34. No. The galvanic cell requires two dissimilar metals. The electrodes are the source of the electrons, not the solutions.

35. a. *Oxidation at anode:* $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

Reduction at cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Overall reaction:

$Mg(s) + 2Ag^+(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$

b. $Cl^-(aq)$ from the NaCl reacts with the $Ag^+(aq)$ to form AgCl(s). This decreases the concentration of the $Ag^+(aq)$. The expected standard conditions no longer apply and the voltage decreases.

36. Al(s), $H_2(g)$, Ag(s), $Cl^-(aq)$

37. The surface of the moon does not have an atmosphere or water. Since oxygen is not available for the redox reaction, iron would not corrode.

38. a. $E°\text{cell} = +1.25\ \text{V}$

Oxidation at anode: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction at cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Overall cell reaction: $2Ag^+(aq) + Fe(s) \rightarrow 2Ag(s) + Fe^{2+}(aq)$

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s) \qquad E°_{\text{anode}} = -\ 0.45\ \text{V}$

$Ag^+(aq) + 1e^- \rightarrow Ag(s) \qquad E°_{\text{cathode}} = +\ 0.80\ \text{V}$

$E°_{\text{cell}} = E°_{\text{cathode}} - E°_{\text{anode}}$

$= +\ 0.80\ \text{V} - (-\ 0.45\ \text{V}) = +\ 1.25\ \text{V}$

b. $E°_{\text{cell}} = -1.25\ \text{V}$ (not spontaneous)

Oxidation at anode: $Ag(s) \rightarrow Ag^+(aq) + e^-$

Reduction at cathode: $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$

Overall cell reaction:

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

c. mass of anode decreases; mass of cathode increases; total mass increases (because 2 mol of Ag are produced for every 1 mol of Fe used)

d. mass of anode decreases; mass of cathode increases; total mass decreases

39. a. anode - inert carbon electrode (negative); cathode - silver electrode (positive)

b. *Oxidation at anode:* $2I^-(aq) \rightarrow I_2(s) + 2e^-$

Reduction at cathode: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$

Overall cell reaction:

$2I^-(aq) + 2Ag^+(aq) \rightarrow I_2(s) + 2Ag(s)$

c. Silver is the oxidizing agent and the iodide ion is the reducing agent.

d. $I_2(s) + 2e^- \rightarrow 2I^-(aq) \qquad E°_{\text{anode}} = 0.54\ \text{V}$

$Ag^+(aq) + 1e^- \rightarrow Zn(s) \qquad E°_{\text{cathode}} = 0.80\ \text{V}$

$E°_{\text{cell}} = E°_{\text{cathode}} - E°_{\text{anode}}$

$= 0.80\ \text{V} - (0.54\ \text{V}) = +0.26\ \text{V}$

40. a. $E°_{\text{cell}} = E°_{\text{cathode}} - E°_{\text{anode}}$

$= -\ 0.42\ \text{V} - (0.82\ \text{V}) = -1.24\ \text{V}$

b. The reaction is not spontaneous as written. Energy must be supply from an external source to make the reaction occur.

41. Ag(s), Cu(s), then Fe(s)

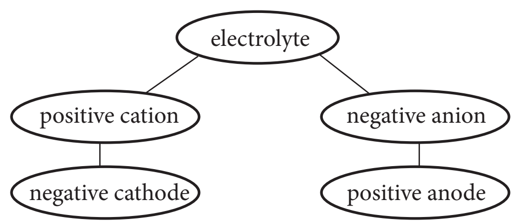
42. The silver ions will diffuse from the salt bridge into the copper (II) nitrate solution over time. Since these silver ions are stronger oxidizing agents than the copper(II) ions, the silver ions will migrate to the copper anode and accept electrons, rather than the copper ions. This would result in an incorrect voltmeter reading for the voltage of the cell of interest in this experiment.

43. Rechargeable batteries can be recycled and should be as they contain a toxic metal (cadmium). An industry funded group called the Rechargeable Battery Recycling Corporation (RBRC) is actively promoting this recycling effort. This corporation is a global corporation that offers a free rechargeable battery and cell phone collection program. Since 1994, Call2Recycle has diverted over 27 million kilograms of rechargeable batteries from landfill sites. In Ontario, companies such as the Home Depot, Staples, Canadian Tire and The Source are drop off sites for the program.

44. Students may say they are against the smelter because the waste products that are formed, such as sulfur dioxide, can contribute to acid precipitation. Students may say they are in favour of the plant because of the economic situation and the fact that the modern mechanisms for removing waste (sulfur dioxide) from the roasting process results in a reduction in the quantities of these pollutants.

Accept all reasoned and reasonable answers, but students should demonstrate some understanding of the processes involved.

45. The identity of the anode and cathode would have to come from knowing the components in the electrolyte. A flow chart is a possibility for a graphic organizer.



46. The anode is a grid made of powdered lead. the cathode is a grid of powdered lead(IV) oxide. The potential of this cell is about 2 V. Six of these are connected in series in an electrolyte of sulfuric acid to give a battery with a potential of 12 V.

47. Answers should reflect the following information:

Galvanic Cell

- Spontaneous reaction
- Converts chemical energy to electrical energy
- Anode (negatively charged): zinc electrode
- Cathode (positively charged): copper electrode
- Oxidation (at anode): $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
- Reduction (at cathode): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
- Cell reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Electrolytic Cell
- Non-spontaneous reaction
- Converts electrical energy to chemical energy
- Anode (positively charged): copper electrode
- Cathode (negatively charged): zinc electrode
- Oxidation (at anode): $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
- Reduction (at cathode): $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
- Cell reaction: $Cu(s) + Zn^{2+}(aq) \rightarrow Cu^{2+}(aq) + Zn(s)$

48. An electrode can be either an anode or a cathode, depending on the other half reaction in the cell. The two half reactions each have a half cell potential that will determine which is the reduction half reaction and which is the oxidation half reaction. The galvanic cell will have a positive cell potential, so the stronger reducing agent will be the anode and the weaker reducing agent will be the cathode.

49. All reports must include a description of the chemistry associated with the occupation. Allow students to present their answer in any of the multiple forms of reporting to allow for an expression of their findings.

50. In whatever form the student chooses here, the student must clearly demonstrate the following

similarities:
- both convert between chemical and electrical energy
- both involve a redox reaction and a cell potential can be calculated
- both involve a reduction at the cathode and an oxidation at the anode

differences:
- galvanic cell has a positively charged cathode and a negatively charged cathode (opposite in an electrochemical cell)
- a galvanic cell is spontaneous while an electrolytic cell in non-spontaneous
- an electrochemical cell requires an external power source to cause the reaction to occur

51. Script should describe the "old" battery technology and its effect on the environment and use chemistry to explain why the changes are an improvement from an environmental point of view.

52. The sketch should look similar to Figure 10.24A and B page 660. Table 10.2 in the textbook summarizes the details of a comparison.

53. After brainstorming, student response could be organized in a PMI organizer. Criteria could include: resources used (cost, accessibility, production safety); production process; waste generated during production; durability; recyclability; disposal process (waste, safety, cost). Details of the criteria must be included.

54. Brochure or other communication piece should include the name and description of the cell technology, how it works (including the chemical processes involved), and a statement or opinion about the viability of the new technology. The submission should also include diagrams or other art showing the technology, and list the sources used for the information.

55. Summary should include all of the key terms and key ideas for the chapter as shown on page 676 of the student textbook.

56. a. *Oxidation:* $2I^-(aq) \rightarrow I_2(s) + 2e^-$

Reduction: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

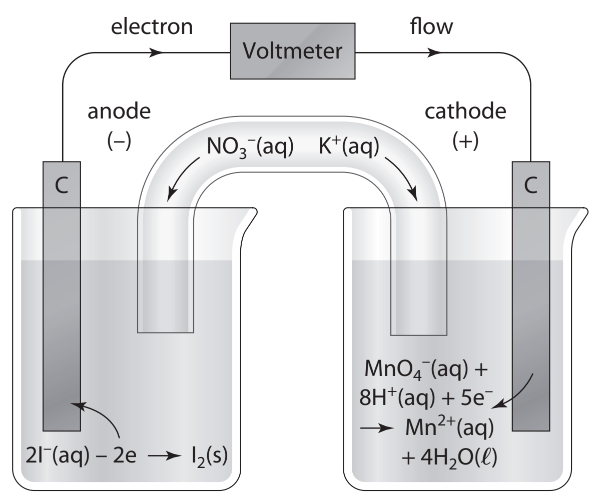
Overall cell reaction:

$10I^-(aq) + 2MnO_4^-(aq) + 16H^+(aq) \rightarrow 5I_2(s) + 2Mn^{2+}(aq) + 8H_2O(\ell)$

b. The oxidizing agent is $MnO_4^-(aq)$ and the reducing agent is $I^-(aq)$.

c. At the anode because it is here that $I^-(aq)$ lose electrons , as reduction occurs

d.



57. $Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $\quad E°_{anode} = -1.66$ V

$Ag^+(aq) + 1e^- \rightarrow Ag(s)$ $\quad E°_{cathode} = 0.80$ V

$E°_{cell} = E°_{cathode} - E°_{anode}$

$= 0.80 - (-1.66\text{ V}) = +2.46\text{ V}$

oxidation, Al/Al^{3+}; reduction, Ag/Ag^+

58. *Oxidation:* $K(s) \rightarrow K^+(aq) + 1e^-$

Reduction: $Fe^{3+}(aq) + 1e^- \rightarrow Fe^{2+}(aq)$

Overall cell reaction:

$K(s) + Fe^{3+}(aq) \rightarrow K^+(aq) + Fe^{2+}(aq)$

59. a. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (anode)

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ (cathode)

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s) \qquad E^\circ_{anode} = -0.76\ V$

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s) \qquad E^\circ_{cathode} = -0.45\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= -0.45\ V - (-0.76\ V) = +0.31\ V$

b. $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$ (anode)

$Al^{3+}(aq) + 3e^- \rightarrow Cr(s)$ (cathode)

$Cr^{3+}(aq) + 3e^- \rightarrow Fe(s) \qquad E^\circ_{anode} = -0.74\ V$

$Al^{3+}(aq) + 3e^- \rightarrow Al(s) \qquad E^\circ_{cathode} = -1.66\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= -1.66 - (-0.74\ V) = -0.92\ V$

c. $2H_2O_2(\ell) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$ (anode)

$Ag^+(aq) + e^- \rightarrow Ag(s)$ (cathode)

$O_2(g) + 2H^+(aq) + 2e^- \rightarrow 2H_2O_2(\ell)$

$E^\circ_{anode} = +0.70\ V$

$Ag^+(aq) + e^- \rightarrow Ag(s) \qquad E^\circ_{cathode} = +0.80\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= 0.80 - (0.70\ V) = +0.10\ V$

60. a. and c. are spontaneous and b. is non-spontaneous

61. a. Estimates should include survey data and method used to arrive at the estimate. Ensure students consider all uses of batteries, including watches, cellular phones, and portable CD players.

b. Students may suggest purchasing rechargeable or longer-life batteries as a partial solution. They could suggest using an AC adapter for battery powered devices whenever possible. Or, students could simply find ways to reduce their use of battery-powered devices.

62. In general, while the chemistry is sound, these devices can only protect a small area surrounding where the device is attached. Car parts tend to be insulated or isolated from each other, and thus the device has no chance of protecting these parts, unless it is directly attached to the part. As well, most car manufacturers now zinc electroplate the entire chassis of the car, which will protect against rust as long as the zinc is present.

63. *net ionic equation (spontaneous)*:

$N_2O(g) + 2H^+(aq) + 2Cu(s) + 2I^-(aq) \rightarrow N_2(g) + H_2O(\ell) + 2CuI(s)$

$E^\circ_{cell} = +1.955\ V$

net ionic equation (non-spontaneous):

$N_2(g) + H_2O(\ell) + 2CuI(s) \rightarrow N_2O(g) + 2H^+(aq) + 2Cu(s) + 2I^-(aq)$

$E^\circ_{cell} = -1.955\ V$

64. a. Yes, because aluminum is reduced more readily than iron. The aluminum would be oxidized, forming a relatively inert oxide coating, and thus hold the iron gutter in place while protecting the iron from oxidation.

b. No, because after the aluminum siding forms the surface oxide, the iron nails would rust and flake to pieces.

65. The battery must be able to supply enough voltage to power the item in which it is being used. The battery must also be able to supply this voltage for as long a period of time as possible. The battery should also be rechargeable and not contain any environmentally harmful chemicals for those batteries that no longer are able to be recharged and thus are disposed of in a landfill site. The size and shape of the battery should also be considered.

66. Diagram and explanation should show one end of the anode of the electroplating cell connected to the positive terminal of an external power supply and the other to a piece of metal that is being used to plate onto the surface of the less expensive metal. This will allow the ions of this metal in solution to be replaced as the plating occurs. The cathode is connected to the negative terminal of the external power supply and then connected to the metal to be plated. As current flows to the cathode, positive metallic ions in solution migrate to the cathode and attach to the surface as they pick up electrons and form a solid surface on the metal.

67. In general, the larger the battery, the more current it can supply, even when the voltages are approximately the same. Therefore, in a device such as a powerful flashlight, a larger current is needed to power the light bulb and so a series of D cell batteries are needed. In a device such as a pen light, where a lower current can be used to light the bulb, AAA batteries are more suited to power the device.

68. In general, the chemistry and design of a disposable battery is such that the cost can be kept to a minimum. The disposable battery tends to be able to hold a charge longer and thus is better in applications where the device is not used as often, or on a low drain item such a TV remote. The design and chemistry of a rechargeable battery needs to be carefully considered to allow for the recharging to occur. This tends to increase the cost of the battery. The chemicals in a rechargeable battery tend to be less environmentally friendly and thus disposal when they can no longer hold a charge is more of an issue.

69. A cell fuelled by aluminum generates power through an electrochemical reaction between the metal, placed in a saline or alkaline solution, and oxygen from the air. Electricity is produced as the aluminum oxidizes (at the anode), so it acts as the reducing agent. Oxygen from the air acts as a cathode. To recharge the cell when the aluminum is consumed, the plates are replaced and more electrolyte is added.

70. In the short term, painting will be effective because it provides a barrier between the metal and moisture and oxygen. However, in the long term, paint can crack and peel and will allow for water and oxygen to oxidize the metal. As well, any object that may scrape against the painted surface can also cause this contact.

71. Cathodic protection using aluminum bars attached to the hull will cause the aluminum be oxidized, rather than the iron in the steel hull. The aluminum is the stronger reducing agent and thus will oxidize. The aluminum is much lighter than a metal such as zinc, and on a floating structure, weight can be a factor that must be considered.

72. Corroding iron mains must be repaired or replaced, which costs money. Corroding mains may also leak and their hydraulic capacity may decrease, which decreases their efficiency. Leaking and corroding water mains may allow contaminants into drinking water. Several ways to combat these problems include replacing iron water mains with polymer or cement mains, which do not corrode, or using sacrificial anodes to prevent corrosion.

73. The battery must be able to supply enough voltage to power the item in which it is being used. The battery must also be able to supply this voltage for as long a period of time as possible. The battery should also be rechargeable and not contain any environmentally harmful chemicals for those batteries that no longer are able to be recharged and thus are disposed of in a landfill site. The size and shape of the battery should also be considered.

74. a. Power tools are large and bulky requiring a large amount of electrical energy over a short burst, e.g., an electric drill. Electronic devices, iPods, cell phones, tablets, etc., require a smaller amount of energy over a longer period of time. These devices are having more apps and are made smaller for easy handling.

b. It is important that the battery be rechargeable and deliver a large amount of electrical energy.

75. If the battery cannot be replaced and recharged, the device would have to be replaced. This is a greater expense for the consumer, but more money for the retailer. It also leads to an increase in electronic waste in the environment. This is a poor idea.

Answers to Chapter 10 Self-Assessment Questions

(Student textbook pages 682-3)

1. e

2. d

3. d

4. c

5. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s) \qquad E° = -0.76\ V$

$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s) \qquad E° = -0.40\ V$

If the reaction is written as
$Zn(s) + Cd^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cd(s)$,
then Zn is the anode and Cd is the cathode.
The answer will be "e" as shown:

$$\begin{aligned} E°_{cell} &= E°_{cathode} - E°_{anode} \\ &= -0.40\ V - 0.76\ V \\ &= +0.36\ V \end{aligned}$$

If the reaction is written as
$Cd(s) + Zn^{2+}(aq) \rightarrow Cd^{2+}(aq) + Zn(s)$,
then Cd is the anode and Zn is the cathode.
The answer will be "d" as shown:

$$\begin{aligned} E°_{cell} &= E°_{cathode} - E°_{anode} \\ &= -0.76\ V - (-0.40)\ V \\ &= -0.36\ V \end{aligned}$$

6. a

7. d

8. c

9. a

10. c

11. The iodide ions will eventually migrate into the lead half-cell and cause lead (II) iodide solid to form. This will remove lead (II) ions from solution and affect the voltage and performance of the galvanic cell.

12. 6

13. Cadmium is toxic and is long-lasting in the environment. In humans, it has been related to high blood pressure and heart disease.

14. a. *Oxidation:* $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Reduction: $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$

Overall cell reaction:

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

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E^\circ_{anode} = -\ 0.76\ V$

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $\quad E^\circ_{cathode} = -\ 0.45\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= -0.45\ V - (-0.76\ V) = +0.31\ V$

b. *Oxidation:* $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Overall cell reaction:

$3Cu^{2+}(aq) + 2Cr(s) \rightarrow 3Cu(s) + 2Cr^{3+}(aq)$

$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ $\quad E^\circ_{anode} = -0.74\ V$

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E^\circ_{cathode} = +0.34\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= 0.34\ V - (-0.74\ V) = +1.08\ V$

c. *Oxidation:* $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Overall cell reaction:

$3Ag^+(aq) + Al(s) \rightarrow 3Ag(s) + Al^{3+}(aq)$

$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $\quad E^\circ_{anode} = -1.66\ V$

$Ag^+(aq) + 1e^- \rightarrow Ag(s)$ $\quad E^\circ_{cathode} = 0.80\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= 0.80 - (-1.66\ V) = +2.46\ V$

15. $K(\ell)$ and $Cl_2(g)$

16. a. *Oxidation:*

$Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(\ell) + 2e^-$

Reduction:

$HgO(s) + H_2O(\ell) + 2e^- \rightarrow Hg(\ell) + 2OH^-(aq)$

b. *Overall cell reaction:*

$Zn(s) + HgO(s) \rightarrow ZnO(s) + Hg(\ell)$

17. The salt added to the roads causes the snow and ice to melt. This melted snow and ice dissolves the salt to form a solution that will be the electrolyte for an electrochemical cell with dissimilar metals or between impurities in the metal. This means that using salt allows for the increase in rust formation on vehicles. Using sand will increase tire traction on snow and ice without leading to an increase in rust formation.

18. a. *Oxidation(anode):* $2I^-(aq) \rightarrow I_2(s) + 2e^-$

Reduction(cathode): $Cl_2(g) + 2e^- \rightarrow Cl^-(aq)$

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $\quad E^\circ_{anode} = +0.54\ V$

$Cl_2(g) + 2e^- \rightarrow Cl^-(aq)$ $\quad E^\circ_{cathode} = 1.36\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= 1.36\ V - (0.54\ V)$

$= +0.82\ V$ spontaneous

b. *Oxidation(anode):*

$2Cr^{3+}(aq) + 7H_2O(\ell) \rightarrow$

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^-$

Reduction(cathode): $Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$

$2Cr^{3+}(aq) + 7H_2O(\ell) + 6e^- \rightarrow$

$2Cr^{3+}(aq) + 7H_2O(\ell)$ $\quad E^\circ_{anode} = 1.36\ V$

$Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$ $\quad E^\circ_{cathode} = 1.07\ V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= 1.07 - (1.36\ V)$

$= -0.29\ V$ non-spontaneous

19. About 95% less energy is required to process used aluminum than is to process the aluminum from the ore in a electrochemical reaction. The aluminum in most products tends to be relatively pure, and there is little surface oxide to remove during the recycling process.

20. a. *Anode:*

$CH_3OH(\ell) + H_2O(\ell) \rightarrow CO_2(g) + 6H^+(aq) + 6e^-$

Cathode: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$

b. Overall cell reaction:

$2CH_3OH(\ell) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(\ell)$

21. a. In the galvanic cell, the anode is the zinc electrode and the cathode is the copper electrode.

b. In an electrolytic cell, the anode is the copper electrode and the cathode is the zinc electrode

22. Fuel cells provide clean electric energy that can be used for transportation, industries and homes. The fuel cell allows for reactants to flow in and products to flow out. The cell converts the energy in the fuel directly into electrical energy without burning the fuel as a combustion engine would. The major waste product of the fuel cell is water.

23. When the steel cans corrode, they produce aqueous ions. However, the aluminum cans produce solid aluminum oxide, which slows down the corrosion process.

24. Iron is abundant (the second most abundant metal after aluminum), relatively inexpensive to mine and process, and can be used to create a number of different alloys with a variety of useful physical properties.

25. $+0.74\ V$

$Zn^{+2}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E^\circ_{anode} = -\ 0.76\ V$

$MnO_2(s) + H_2O(\ell) + 2e^- \rightarrow$

$Mn_2O_3(s) + 2OH^-(aq)$ $\quad E^\circ_{cathode} = X$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$1.50\ V = X - (-0.76\ V)$

$X = +\ 0.74\ V$

Answers to Unit Review Questions

(Student textbook pages 687-91)

1. a

2. c

3. e

4. b

5. c

6. e

7. b

8. d

9. c

10. b

11. a. +4

b. +7

c. +5

d. +3

12. A primary battery cannot be recharged, while a secondary battery can be recharged.

13. An inert electrode is made from a material that is neither a reactant nor a product of the redox reaction but can carry a current and provide a surface on which redox reactions can occur. Carbon and platinum are commonly used as inert electrodes.

14. "Standard reduction potential" refers to the amount of energy conferred on a unit of charge as the material listed is reduced. Since potential must be measured in reference to another half-cell, the hydrogen half-cell is used as the reference point; its reduction potential is set at 0.00 V and all other half-cells are measured relative to this.

15. a. Metal A is the more effective reducing agent. It is oxidized and facilitates the reduction of the other metal. For example, in the following reaction, magnesium is the more effective reducing agent.

$Mg(s) + Zn(NO_3)_2(aq) \rightarrow Zn(s) + Mg(NO_3)_2(aq)$

b. Non-metal A is the more effective oxidizing agent. For one non-metal to replace another during a single displacement reaction, it must gain electrons more easily than the non-metal it is displacing. Therefore, it must be a better oxidizing agent. An example is:

$F_2(g) + 2NaCl(aq) \rightarrow 2NaF(aq) + Cl_2(g)$

16. A balanced half-reaction equation indicates whether oxidation or reduction is taking place, how many electrons are involved, and identifies either the oxidizing agent or the reducing agent. For example, consider the following equations:

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Because the electrons are on the products side, the half-reaction is an oxidation. Zinc is a reducing agent. Each zinc atom loses two electrons, and Zn^{2+} is produced.

$Cu^{2+} + 2e^- \rightarrow Cu(s)$

Because the electrons are on the reactants side, the half-reaction is a reduction. Copper(II) ions are an oxidizing agent. Each copper ion gains two electrons, and copper metal is produced.

17. A more electronegative material will tend to pull electrons towards it, therefore gaining electrons meaning that it will be reduced. An atom of lower electronegativity will not hold electrons as tightly, and will lose electrons and thus be oxidized.

18. Oxygen atoms are not always involved in the redox process. For example, in a half-reaction such as $NO^{3-}(aq) \rightarrow NO^{2-}(aq)$, the oxygen atoms have an oxidation number of −2, which means that they do not go through oxidation or reduction in the process. In addition, the oxygen atoms are in the compound that is reduced as well as the compound that is oxidized.

19. a. Because oxidation occurs at the anode, the gas must be oxygen gas.

b. A glowing splint placed into a container of this gas would burst into flames if the gas is oxygen.

20. The definition had to be changed because chemists started to recognize similarities between reactions of atoms and compounds with oxygen and reactions of the same atoms and compounds with elements other than oxygen. The change in the elements reacting with other elements was identical to the change in the same elements when they reacted with oxygen.

21. Elemental carbon is oxidized to carbon monoxide in the reaction: $C(s) + O_2(g) \rightarrow CO(g)$

The carbon monoxide reduces iron in several reactions.

$3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$

$Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$

$FeO(s) + CO(g) \rightarrow Fe(\ell) + CO_2(g)$

22. A redox reaction where atoms of one element undergo both oxidation and reduction in a single reaction.

a. $2Cu^+(aq) \rightarrow Cu(s) + Cu^{2+}(aq)$

b. *Oxidation:* $Cu^+(aq) \rightarrow Cu^{2+}(aq) + e^-$

Reduction: $Cu^+(aq) + e^- \rightarrow Cu(s)$

23. The oxidation number that is calculated according to the rules is simply the average oxidation number for all of the carbon atoms in the molecule. In cases such as this, each carbon atom in the same molecule has its own unique oxidation number. In $H_6C_3O(\ell)$, the oxidation numbers of the three carbon atoms are −2, −3, and +1. The average is $-\frac{4}{3}$.

24. When the reactant and product of a half reaction that is occurring in one cell are both in solution, a surface is needed on which the reaction can occur. This inert electrode is the location for electron transfer as the half-reaction occurs. The inert electrode also carries the electron flow to the external circuit of the cell.

25. *Step 1:* Balance hydrogen atoms by adding hydrogen ions.

Step 2: Add the same number of hydroxide ions to both sides as the number of hydrogen ions added in step 1.

Step 3: On the side with hydrogen ions and hydroxide ions, combine to form water molecules, and then cancel water molecules present on both sides of the equation.

Step 4: Balance charges by adding electrons.

26. *Sample answer*: For the anode of the galvanic cell, I would use lead in a solution of lead sulfate. For the cathode, I would use silver in a solution of silver sulfate. A power source changes the galvanic cell to an electrolytic cell. For the cathode of the electrolytic cell, I would use lead in a solution of lead sulfate. For the anode, I would use silver in a solution of silver sulfate.

27. a. $8H^+(aq) + MnO^{4-}(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

b. $SO_3^{2-}(aq) + 2OH^-(aq) \rightarrow SO_4^{2-}(aq) + H_2O(\ell) + 2e^-$

c. $NO_3^-(aq) + 7H_2O(\ell) + 8e^- \rightarrow NH_4^+(aq) + 10OH^-(aq)$

d. $Re^-(aq) + 3H_2O(\ell) \rightarrow ReO_3^-(aq) + 6H^+(aq) + 6e^-$

28. a. reduction

b. oxidation

c. reduction

d. oxidation

29. a. *oxidation*: $Cd(s) \rightarrow Cd^{2+}(aq)$

$Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$

reduction:

$NO_3^-(aq) \rightarrow NO(g)$

$NO_3^-(aq) \rightarrow NO(g) + 2H_2O(\ell)$

$NO_3^-(aq) + 4H^+(aq) \rightarrow NO(g) + 2H_2O(\ell)$

$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

multiply oxidation rx. by 3:

$3[Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-]$

$3Cd(s) \rightarrow 3Cd^{2+}(aq) + 6e^-$

multiply reduction rx. by 2:

$2[NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)]$

$2NO_3^-(aq) + 8H^+(aq) + 6e^- \rightarrow 2NO(g) + 4H_2O(\ell)$

combine:

$3Cd(s) + 2NO_3^-(aq) + 8H^+(aq) \cancel{+ 6e^-} \rightarrow 3Cd^{2+}(aq) \cancel{+ 6e^-} + 2NO(g) + 4H_2O(\ell)$

balance:

$3Cd(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cd^{2+}(aq) + 2NO(g) + 4H_2O(\ell)$

b. (This problem is much more easily done by inspection. However, the directions say to use the half-reaction method. Because hydroxide ions are involved, it will be solved in a basic solution.)

oxidation: $Fe^{2+} \rightarrow Fe^{3+}$

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

reduction:

$2OH^-(aq) + H_2O_2(\ell) \rightarrow 3OH^-(aq)$

$2OH^-(aq) + H_2O_2(\ell) \rightarrow 3OH^-(aq) + H_2O(\ell)$

$2OH^-(aq) + H_2O_2(\ell) + H^+(aq) \rightarrow 3OH^-(aq) + H_2O(\ell)$

$2OH^-(aq) + H_2O_2(\ell) + H^+(aq) + OH^-(aq) \rightarrow 3OH^-(aq) + H_2O(\ell) + OH^-(aq)$

$2OH^-(aq) + H_2O_2(\ell) + H_2O(\ell) + 2e^- \rightarrow 4OH^-(aq) + H_2O(\ell)$

multiply oxidation rx. by 2: $2[Fe^{2+} \rightarrow Fe^{3+} + e^-]$

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

multiply reduction rx. by 1:

$2OH^-(aq) + H_2O_2(\ell) + H_2O(\ell) + 2e^- \rightarrow 4OH^-(aq) + H_2O(\ell)$

combine:

$2Fe^{2+} + 2OH^-(aq) + H_2O_2(\ell) + H_2O(\ell) + 2e^- \rightarrow 2Fe^{3+} + 2e^- + 4OH^-(aq) + H_2O(\ell)$

$2Fe^{2+} + 2OH^-(aq) + H_2O_2(\ell) + \cancel{H_2O(\ell) + 2e^-} \rightarrow 2Fe^{3+} \cancel{+ 2e^-} + 4OH^-(aq) + \cancel{H_2O(\ell)}$

balance:

$2Fe^{2+} + 2OH^-(aq) + H_2O_2(\ell) \rightarrow 2Fe^{3+} + 4OH^-(aq)$

$2Fe^{2+} + 2OH^-(aq) + H_2O_2(\ell) + 2OH^-(aq) \rightarrow 2Fe^{3+} + 4OH^-(aq) + 2OH^-(aq)$

$2Fe(OH)_2(aq) + H_2O_2(\ell) \rightarrow 2Fe(OH)_3(aq)$

c. oxidation: $I_2(s) \rightarrow IO_3^-(aq)$

$I_2(s) \rightarrow 2IO_3^-(aq)$

$I_2(s) + 6H_2O(\ell) \rightarrow 2IO_3^-(aq)$

$I_2(s) + 6H_2O(\ell) \rightarrow 2IO_3^-(aq) + 12H^+(aq)$

$I_2(s) + 6H_2O(\ell) + 12OH^-(aq) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 12OH^-(aq)$

$I_2(s) + 6H_2O(\ell) + 12OH^-(aq) \rightarrow 2IO_3^-(aq) + 12H_2O(\ell)$

$I_2(s) + 6H_2O(\ell) + 12OH^-(aq) \rightarrow 2IO_3^-(aq) + 12H_2O(\ell) + 10e^-$

reduction: $I_2(s) \rightarrow I^-(aq)$

$I_2(s) \rightarrow 2I^-(aq)$

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$

multiply oxidation rx. by 1:

$I_2(s) + 6H_2O(\ell) + 12OH^-(aq) \rightarrow 2IO_3^-(aq) + 12H_2O(\ell) + 10e^-$

multiply reduction rx. by 5:

$5I_2(s) + 10e^- \rightarrow 10I^-(aq)$

combine:

$I_2(s) + 6H_2O(\ell) + 12OH^-(aq) + 5I_2(s) \cancel{+ 10e^-} \rightarrow 2IO_3^-(aq) + 12H_2O(\ell) \cancel{+ 10e^-} + 10I^-(aq)$

balance: $I_2(s) + \cancel{6H_2O(\ell)} + 12OH^-(aq) + 5I_2(s) \rightarrow 2IO_3^-(aq) + \cancel{12}H_2O(\ell) + 10I^-(aq)$

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

$3I_2(s) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(\ell) + 5I^-(aq)$

30. a. Assign oxidation numbers:

$$\underset{+3}{Cr^{3+}}(aq) + \underset{+5\ -2}{ClO_3^-}(aq) \longrightarrow \underset{+6\ -2}{Cr_2O_7^{2-}}(aq) + \underset{+4\ -2}{ClO_2}(aq)$$

Track electrons:

1(-3e⁻) = -3e⁻

$$Cr^{3+}(aq) + ClO_3^-(aq) \longrightarrow Cr_2O_7^{2-}(aq) + ClO_2(aq)$$

3(+1e⁻) = +3e⁻

Place coefficients determined by the gain and loss of electrons:

$Cr^{3+}(aq) + 3ClO_3^-(aq) \rightarrow Cr_2O_7^{2-}(aq) + ClO_2(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the chromium and chlorine atoms remain as above:

$2Cr^{3+}(aq) + 6ClO_3^-(aq) \rightarrow Cr_2O_7^{2-}(aq) + 6ClO_2(aq)$

Add water to balance the oxygen atoms:

$2Cr^{3+}(aq) + 6ClO_3^-(aq) + H_2O(\ell) \rightarrow Cr_2O_7^{2-}(aq) + 6ClO_2(aq)$

Add hydrogen ions to balance the hydrogen atoms:

$2Cr^{3+}(aq) + 6ClO_3^-(aq) + H_2O(\ell) \rightarrow Cr_2O_7^{2-}(aq) + 6ClO_2 + 2H^+(aq)$

b. Assign oxidation numbers:

$$\underset{-1}{I^-}(aq) + \underset{+7\ -2}{MnO_4^-}(aq) \longrightarrow \underset{0}{I_2}(s) + \underset{+2}{Mn^{2+}}(aq)$$

Track electrons:

5(-1e⁻) = -5e⁻

$$I^-(aq) + MnO_4^-(aq) \longrightarrow I_2(s) + Mn^{2+}(aq)$$

1(+5e⁻) = +5e⁻

Place coefficients determined by the gain and loss of electrons:

$5I^-(aq) + MnO_4(aq) \rightarrow I_2(s) + Mn^{2+}(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the iodine and manganese atoms remain as above:

$10I^-(aq) + 2MnO_4^-(aq) \rightarrow 5I_2(s) + 2Mn^{2+}(aq)$

Add water to balance the oxygen atoms:

$10I^-(aq) + 2MnO_4^-(aq) \rightarrow 5I_2(s) + 2Mn^{2+}(aq) + 8H_2O(\ell)$

Add hydrogen ions to balance the hydrogen atoms:

$10I^-(aq) + 2MnO_4^-(aq) + 16H^+(aq) \rightarrow 5I_2(s) + 2Mn^{2+}(aq) + 8H_2O(\ell)$

c. Assign oxidation numbers:

$$\underset{+2\ -3}{CN^-}(aq) + \underset{+2\ \ -3\ +1}{Cu(NH_3)_4^{2+}}(aq) \longrightarrow \underset{+1\ \ +2\ -3}{Cu(CN)_3^{2-}}(aq) + \underset{-3\ +1}{NH_3}(g) + \underset{+4\ -3\ -2}{CNO^-}(aq)$$

Track electrons:

1(-2e⁻) = -2e⁻

$$CN^-(aq) + Cu(NH_3)_4^{2+}(aq) \longrightarrow Cu(CN)_3^{2-}(aq) + NH_3(g) + CNO^-(aq)$$

2(+1e⁻) = +2e⁻

Place coefficients determined by the gain and loss of electrons:

$CN^-(aq) + 2Cu(NH_3)_4^{2+}(aq) \rightarrow 2Cu(CN)_3^{2-}(aq) + NH_3(g) + CNO^-(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the carbon and copper atoms remain as above: (Notice that the CN^- in the $Cu(CN)_3^{2-}$ do not change in oxidation number. To help keep them separate from the CN^- in which the carbon changes oxidation number, add another set of CN^- ions to the left side to balance those in the $Cu(CN)_3^{2-}$.)

$CN^-(aq) + 2Cu(NH_3)_4^{2+}(aq) + 6CN^-(aq) \rightarrow 2Cu(CN)_3^{2-}(aq) + 8NH_3(g) + CNO^-(aq)$

Add water to balance the oxygen atoms:

$CN^-(aq) + 2Cu(NH_3)_4{}^{2+}(aq) + 6CN^-(aq) + H_2O(\ell)$
$\rightarrow 2Cu(CN)_3{}^{2-}(aq) + 8NH_3(g) + CNO^-(aq)$

Add hydrogen ions to balance the hydrogen atoms:

$CN^-(aq) + 2Cu(NH_3)_4{}^{2+}(aq) + 6CN^-(aq) + H_2O(\ell)$
$\rightarrow 2Cu(CN)_3{}^{2-}(aq) + 8NH_3(g) + CNO^-(aq) + 2H^+(aq)$

To both sides, add the number of hydroxide ions equal to the number of hydrogen ions.

$CN^-(aq) + 2Cu(NH_3)_4{}^{2+}(aq) + 6CN^-(aq) + H_2O(\ell) + 2OH^-(aq) \rightarrow$
$2Cu(CN)_3{}^{2-}(aq) + 8NH_3(g) + CNO^-(aq) + 2H^+(aq) + 2OH^-(aq)$

Combine H^+ and OH^- to make water, where possible.

$CN^-(aq) + 2Cu(NH_3)_4{}^{2+}(aq) + 6CN^-(aq)$ $\cancel{+ H_2O(\ell)} + 2OH^-(aq) \rightarrow$
$2Cu(CN)_3{}^{2-}(aq) + 8NH_3(g) + CNO^-(aq) + \cancel{2}H_2O(\ell)$

Cancel water molecules that are on both sides of the equation and check to ensure that atoms of all elements are balanced. Combine the CN^- ions at this last step.

$7CN^-(aq) + 2Cu(NH_3)_4{}^{2+}(aq) + 2OH^-(aq) \rightarrow$
$2Cu(CN)_3{}^{2-}(aq) + 8NH_3(g) + CNO^-(aq) + H_2O(\ell)$

31. a. $2Fe^{2+}(aq) + I_2(s) \rightarrow 2Fe^{3+}(aq) + 2I^-(aq)$

$2Fe^{3+}(aq) + 2e^- \rightarrow 2Fe^{2+}$ $\quad E°_{anode} = +0.77\ V$

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $\quad E°_{cathode} = +0.54\ V$

$E°_{cell} = E°_{cathode} - E°_{anode}$
$= +0.54\ V - 0.77\ V$
$= -0.23\ V$

b. $Au(NO_3)_3(aq) + 3Ag(s) \rightarrow 3AgNO_3(aq) + Au(s)$

$3Ag^+(aq) + 3e^- \rightarrow 3Ag(s)$ $\quad E°_{anode} = +0.80\ V$

$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$ $\quad E°_{cathode} = +1.50\ V$

$E°_{cell} = E°_{cathode} - E°_{anode}$
$= +1.50\ V - 0.80\ V$
$= -0.70\ V$

c. $H_2O_2(aq) + 2HCl(aq) \rightarrow Cl_2(g) + 2H_2O(\ell)$

$Cl_2(g) + 2e^- \rightarrow 2Cl^-(g)$ $\quad E°_{anode} = +1.36\ V$

$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O\ (\ell)$ $\quad E°_{cathode} = +1.78\ V$

$E°_{cell} = E°_{cathode} - E°_{anode}$
$= +1.78\ V - 1.36\ V$
$= +0.42\ V$

32. a. non-spontaneous

b. spontaneous

c. spontaneous

33. *Ionic equation:*

$Cd(s) + Sn^{2+}(aq) + 2NO_3{}^-(aq) \rightarrow$
$Cd^{2+}(aq) + 2NO_3{}^-(aq) + Sn(s)$

Net ionic equation:

$Cd(s) + Sn^{2+}(aq) \rightarrow Cd^{2+}(aq) + Sn(s)$

a. A spectator ion does not undergo any changes and is found on both sides of the equation in the same form. The nitrate ion is the spectator ion in the above reaction. It is removed when the net ionic equation is written.

b. Cd(s)

c. $Sn^{2+}(aq)$

d. cadmium

e. tin

34. Student procedures should include an indicator of how they plan on identifying the anode and cathode. For example, they can suggest allowing the reaction to proceed, and then later measure any change in mass of the electrodes. The electrode that gains mass is the cathode, and the electrode that loses mass is the anode. Alternatively, they could use a galvanometer to observe the direction of the current.

35. Half-reactions represent oxidations or reductions, which never occur independently. They are written so they have the smallest possible coefficients, so the number of electrons lost in the oxidation half-reaction might not be the same as the number of electrons gained in the reduction half-reaction. However, when combining the half reactions, the electrons must be balanced. For each electron "lost" in an oxidation reaction, an electron must be "gained" in a reduction reaction. Electrons never appear in the overall equation for a redox reaction, which is consistent with the fact that electrons are not created or destroyed in chemical reactions.

36. a. $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$

b. The reaction is not a redox reaction because the oxidation numbers do not change.

37. Students could choose to construct an electrolytic cell with an iron nail at the cathode in a solution of any cation listed below it on the reactivity series (such as zinc or aluminum).

38. a. $Fe(s) + HNO_3(aq) \rightarrow Fe^{2+}(aq) + NO(g)$

Use the half-reaction method under acidic conditions.

oxidation:

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

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

reduction:
$NO_3^-(aq) \rightarrow NO(g)$
$NO_3^-(aq) \rightarrow NO(g) + 2H_2O(\ell)$
$NO_3^-(aq) + 4H^+(aq) \rightarrow NO(g) + 2H_2O(\ell)$
$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

multiply oxidation rx. by 3:
$3[Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-]$
$3Fe(s) \rightarrow 3\ Fe^{2+}(aq) + 6e^-$

multiply reduction rx. by 2:
$2[NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)]$
$2NO_3^-(aq) + 8H^+(aq) + 6e^- \rightarrow 2NO(g) + 4H_2O(\ell)$

combine:
$3Fe(s) + 2NO_3^-(aq) + 8H^+(aq) \cancel{+ 6e^-} \rightarrow 3\ Fe^{2+}(aq) \cancel{+ 6e^-} + 2NO(g) + 4H_2O(\ell)$

balance:
$3Fe(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Fe^{2+}(aq) + 2NO(g) + 4H_2O(\ell)$

Because you know that the reaction takes place in nitric acid, you can add enough nitrate ions to neutralize the charge and write the complete balanced equation.
$3Fe(s) + 8HNO_3(aq) \rightarrow 3Fe(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$

b. $Fe(NO_3)_2(aq) + HNO_3(aq) \rightarrow Fe(NO_3)_3(aq) + NO(g)$

Use the half-reaction method under acidic conditions.

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

reduction:
$NO_3^- \rightarrow NO(g)$
$NO_3^- \rightarrow NO(g) + 2H_2O(\ell)$
$NO_3^- + 4H^+(aq) \rightarrow NO(g) + 2H_2O(\ell)$
$NO_3^- + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

multiply oxidation rx. by 3:
$3[Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-]$
$3Fe^{2+}(aq) \rightarrow 3Fe^{3+}(aq) + 3e^-$

multiply reduction rx. by 1:
$NO_3^- + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

combine:
$3Fe^{2+}(aq) + NO_3^- + 4H^+(aq) \cancel{+ 3e^-} \rightarrow 3Fe^{3+}(aq) \cancel{+ 3e^-} + NO(g) + 2H_2O(\ell)$

balance:
$3Fe^{2+}(aq) + NO_3^- + 4H^+(aq) \rightarrow 3Fe^{3+}(aq) + NO(g) + 2H_2O(\ell)$

Because you know that the reaction takes place in nitric acid, you can add enough nitrate ions to neutralize the charge and write the complete balanced equation.
$3Fe(NO_3)_2(aq) + 4HNO_3(aq) \rightarrow 3Fe(NO_3)_3(aq) + NO(g) + 2H_2O(\ell)$

c. $Fe(s) + HNO_3(aq) \rightarrow Fe^{3+}(aq) + NO(g)$

oxidation:
$Fe(s) \rightarrow Fe^{3+}(aq)$
$Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$

reduction:
$NO_3^-(aq) \rightarrow NO(g)$
$NO_3^-(aq) \rightarrow NO(g) + 2H_2O(\ell)$
$NO_3^-(aq) + 4H^+(aq) \rightarrow NO(g) + 2H_2O(\ell)$
$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

multiply oxidation rx. by 1:
$Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$

multiply reduction rx. by 1:
$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

combine:
$Fe(s) + NO_3^-(aq) + 4H^+(aq) \cancel{+ 3e^-} \rightarrow Fe^{3+}(aq) \cancel{+ 3e^-} + NO(g) + 2H_2O(\ell)$

balance:
$Fe(s) + NO_3^-(aq) + 4H^+(aq) \rightarrow Fe^{3+}(aq) + NO(g) + 2H_2O(\ell)$

Because you know that the reaction takes place in nitric acid, you can add enough nitrate ions to neutralize the charge and write the complete balanced equation.
$Fe(s) + 4HNO_3(aq) \rightarrow Fe(NO_3)_3(aq) + NO(g) + 2H_2O(\ell)$

d. These all represent corrosion reactions.

39. a. *oxidation*:
$Zn(s) \rightarrow Zn^{2+}(aq)$
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

reduction:
$MnO_2(s) \rightarrow Mn_2O_3(s)$
$2MnO_2(s) \rightarrow Mn_2O_3(s)$
$2MnO_2(s) \rightarrow Mn_2O_3(s) + H_2O(\ell)$
$2MnO_2(s) + 2H^+(aq) \rightarrow Mn_2O_3(s) + H_2O(\ell)$
$2MnO_2(s) + 2H^+(aq) + 2OH^-(aq) \rightarrow Mn_2O_3(s) + H_2O(\ell) + 2OH^-(aq)$
$2MnO_2(s) + 2H_2O(\ell) \rightarrow Mn_2O_3(s) + H_2O(\ell) + 2OH^-(aq)$
$2MnO_2(s) + 2H_2O(\ell) + 2e^- \rightarrow Mn_2O_3(s) + H_2O(\ell) + 2OH^-(aq)$

multiply oxidation rx. by 1:
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

multiply reduction rx. by 1:
$2MnO_2(s) + 2H_2O(\ell) + 2e^- \rightarrow Mn_2O_3(s) + H_2O(\ell) + 2OH^-(aq)$

combine:

$Zn(s) + 2MnO_2(s) + \cancel{2}H_2O(\ell) + \cancel{2e^-} \rightarrow$
$Zn^{2+}(aq) + \cancel{2e^-} + Mn_2O_3(s) + \cancel{H_2O(\ell)} + 2OH^-(aq)$

balance:

$Zn(s) + 2MnO_2(s) + H_2O(\ell) \rightarrow$
$Zn^{2+}(aq) + Mn_2O_3(s) + 2OH^-(aq)$

b. Anode: Zn(s); cathode: graphite rod

40. a. $Cl_2O_7(aq) + H_2O(\ell) \rightarrow 2HClO_4(aq)$

Assign oxidation numbers:

$$\underset{+7\ -2}{Cl_2O_7}(\ell) + \underset{+1\ -2}{H_2O}(\ell) \longrightarrow 2\underset{+1\ +7\ -2}{HClO}\ (aq)$$

The oxidation numbers do not change; therefore, this is not a redox reaction.

b. Assign oxidation numbers:

$$\underset{0}{I_2}(s) + \underset{+5\ -2}{ClO_3^-}(aq) \longrightarrow \underset{+5\ -2}{IO_3^-}(aq) + \underset{-1}{Cl^-}(aq)$$

The oxidation number for iodine increases and the oxidation number for chlorine decreases, so this is a redox reaction.

oxidizing agent: $ClO_3^-(aq)$;

reducing agent: $I_2(s)$

oxidation:

$I_2(s) \rightarrow IO_3^-(aq)$

$I_2(s) \rightarrow 2IO_3^-(aq)$

$I_2(s) + 6H_2O(\ell) \rightarrow 2IO_3^-(aq)$

$I_2(s) + 6H_2O(\ell) \rightarrow 2IO_3^-(aq) + 12H^+(aq)$

$I_2(s) + 6H_2O(\ell) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^-$

reduction:

$ClO_3^-(aq) \rightarrow Cl^-(aq)$

$ClO_3^-(aq) \rightarrow Cl^-(aq) + 3H_2O(\ell)$

$ClO_3^-(aq) + 6H^+(aq) \rightarrow Cl^-(aq) + 3H_2O(\ell)$

$ClO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow Cl^-(aq) + 3H_2O(\ell)$

multiply oxidation rx. by 3:

$3[I_2(s) + 6H_2O(\ell) \rightarrow 2IO_3^-(aq) + 6H^+(aq) + 10e^-]$

$3I_2(s) + 18H_2O(\ell) \rightarrow$
$6IO_3^-(aq) + 18H^+(aq) + 30e^-$

multiply reduction rx. by 5:

$5[ClO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow$
$Cl^-(aq) + 3H_2O(\ell)]$

$5ClO_3^-(aq) + 30H^+(aq) + 30e^- \rightarrow$
$5Cl^-(aq) + 15H_2O(\ell)$

combine:

$3I_2(s) + \cancel{18}H_2O(\ell) + 5ClO_3^-(aq) + 30H^+(aq) + \cancel{30e^-}$
$\rightarrow 6IO_3^-(aq) + 36H^+(aq) + \cancel{30e^-} + 5Cl^-(aq) + \cancel{15H_2O(\ell)}$

balance:

$3I_2(s) + 3H_2O(\ell) + 5ClO_3^-(aq) \rightarrow$
$6IO_3^-(aq) + 6H^+(aq) + 5Cl^-(aq)$

c. Assign oxidation numbers:

$$\underset{-2}{S^{2-}}(aq) + \underset{0}{Br_2}(\ell) \longrightarrow \underset{+6\ -2}{SO_4^{2-}}(aq) + \underset{-1}{Br^-}(aq)$$

The oxidation numbers changed, so this is a redox reaction.

Sulfur's oxidation number went from −2 to +6, so it was oxidized; thus $S^{2-}(aq)$ is the reducing agent.

Bromine's oxidation number went from 0 to −1, so it was reduced; thus $Br_2(g)$ is the oxidizing agent.

oxidation:

$S^{2-}(aq) \rightarrow SO_4^{2-}(aq)$

$S^{2-}(aq) + 4H_2O(\ell) \rightarrow SO_4^{2-}(aq)$

$S^{2-}(aq) + 4H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 8H^+(aq)$

$S^{2-}(aq) + 4H_2O(\ell) + 8OH^-(aq) \rightarrow$
$SO_4^{2-}(aq) + 8H^+(aq) + 8OH^-(aq)$

$S^{2-}(aq) + 4H_2O(\ell) + 8OH^-(aq) \rightarrow$
$SO_4^{2-}(aq) + 8H_2O(\ell)$

$S^{2-}(aq) + 4H_2O(\ell) + 8OH^-(aq) \rightarrow$
$SO_4^{2-}(aq) + 8H_2O(\ell) + 8e^-$

reduction: $Br_2(\ell) \rightarrow Br^-(aq)$

$Br_2(\ell) \rightarrow 2Br^-(aq)$

$Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$

multiply oxidation rx. by 1:

$S^{2-}(aq) + 4H_2O(\ell) + 8OH^-(aq) \rightarrow$
$SO_4^{2-}(aq) + 8H_2O(\ell) + 8e^-$

multiply reduction rx. by 4:

$4[Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)]$

$4Br_2(\ell) + 8e^- \rightarrow 8Br^-(aq)$

combine:

$S^{2-}(aq) + \cancel{4H_2O(\ell)} + 8OH^-(aq) + 4Br_2(\ell) + \cancel{8e^-}$
$\rightarrow SO_4^{2-}(aq) + \cancel{8}H_2O(\ell) + \cancel{8e^-} + 8Br^-(aq)$

balance:

$S^{2-}(aq) + 8OH^-(aq) + 4Br_2(\ell) \rightarrow$
$SO_4^{2-}(aq) + 4H_2O(\ell) + 8Br^-(aq)$

d. Assign oxidation numbers:

$$\underset{+1\ +5\ -2}{HNO_3}(aq) + \underset{+1\ -2}{H_2S}(g) \longrightarrow \underset{+2\ -2}{NO}(g) + \underset{0}{S}(s) + \underset{+1\ -2}{H_2O}(\ell)$$

The oxidation numbers changed, so this is a redox reaction.

Nitrogen's oxidation number went from +5 to +2, so it was reduced; thus $HNO_3(aq)$ is the oxidizing agent.

Sulfur's oxidation number went from −1 to 0, so it was oxidized; thus $H_2S(g)$ is the reducing agent.

Track electrons:

$2(+3e^-) = +6e^-$

$$HNO_3(aq) + H_2S(g) \longrightarrow NO(g) + S(s) + H_2O(\ell)$$

$3(-2e^-) = -6e^-$

Place coefficients determined by the gain and loss of electrons:

$2HNO_3(aq) + 3H_2S(g) \rightarrow NO(g) + S(s) + H_2O(\ell)$

Because this is a neutral solution, you can now balance everything while ensuring that the ratios of the nitrogen and sulfur atoms remain as above:

$2HNO_3(aq) + 3H_2S(g) \rightarrow$
$2NO(g) + 3S(s) + 4H_2O(\ell)$

41. a. *oxidizing agent*: $Ag^+(aq)$; *reducing agent*: $Cu(s)$

b. *oxidizing agent*: the iodine in $IO_3^-(aq)$; *reducing agent*: $Mg(s)$

42. a. spontaneous

b. spontaneous

43. a. Because the reaction was spontaneous, element A must be a stronger reducing agent.

b. B would most likely have the larger electronegativity.

44. a. *oxidation numbers*: The C_2H_5OH does not ionize so it is not feasible to write half reactions.

b. *oxidizing agent*: the chromium in $Cr_2O_7^{2-}$; r*educing agent*: the carbon in the ethanol

c. Assign oxidation numbers:

$$\underset{+6\ \ -2}{Cr_2O_7^{2-}}(aq) + \underset{+2\ +1\ -2\ +1}{C_2H_5OH}(\ell) \longrightarrow \underset{+3}{Cr^{3+}}(aq) + \underset{+4\ -2}{CO_2}(g) + \underset{+1\ \ -2}{H_2O}$$

Track electrons:

$$\overbrace{Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) \longrightarrow Cr^{3+}}^{2(+6e^-) = +12e^-}(aq) + CO_2(g) + H_2O(\ell)$$

$$Cr_2O_7^{2-}(aq) + \underbrace{C_2H_5OH(\ell) \longrightarrow Cr^{3+}(aq) + CO_2}_{1(-12e^-) = -12e^-}(g) + H_2O(\ell)$$

Notice that there are two chromium atoms each gaining three electrons for a total of six, and two carbon atoms each losing six electrons for a total of twelve. Place coefficients determined by the gain and loss of electrons:

$2Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) \rightarrow Cr^{3+}(aq) + CO_2(g)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the chromium and carbon atoms remain as above:

$2Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) \rightarrow$
$4Cr^{3+}(aq) + 2CO_2(g)$

Add water to balance the oxygen atoms:

$2Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) \rightarrow$
$4Cr^{3+}(aq) + 2CO_2(g) + 11H_2O(\ell)$

Add hydrogen ions to balance the hydrogen atoms:

$2Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) + 16H^+(aq) \rightarrow$
$4Cr^{3+}(aq) + 2CO_2(g) + 11H_2O(\ell)$

To both sides, add the number of hydroxide ions equal to the number of hydrogen ions.

$2Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) + 16H^+(aq) + 16OH^-(aq) \rightarrow$
$4Cr^{3+}(aq) + 2CO_2(g) + 11H_2O(\ell) + 16OH^-(aq)$

Combine H^+ and OH^- to make water, where possible.

$2Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) + 16H_2O(\ell) \rightarrow$
$4Cr^{3+}(aq) + 2CO_2(g) + 11H_2O(\ell) + 16OH^-(aq)$

Cancel water molecules that are on both sides of the equation and check to ensure that atoms of all elements are balanced.

$2Cr_2O_7^{2-}(aq) + C_2H_5OH(\ell) + 5H_2O(\ell) \rightarrow$
$4Cr^{3+}(aq) + 2CO_2(g) + 16OH^-(aq)$

45. *Cathode:* $K^+(\ell) + e^- \rightarrow K(\ell)$

Anode: $2F^-(\ell) \rightarrow F_2(g) + 2e^-$

Overall reaction: $2K^+(\ell) + 2F^-(\ell) \rightarrow 2K(\ell) + F_2(g)$

46. *Electrolyte*: $H_2SO_4(aq)$; *anode*: $Pb(s)$; *cathode*: $PbO_2(s)$

47. Assign oxidation numbers:

$$\underset{+7\ \ -2}{MnO_4^-}(aq) + \underset{-3\ \ -2}{C_2O_4^{2-}}(aq) \longrightarrow \underset{+2}{Mn^{2+}}(aq) + \underset{+4\ -2}{CO_2}(g)$$

Track electrons:

$$\overbrace{MnO_4^-(aq) + C_2O_4^{2-}(aq) \longrightarrow Mn^{2+}}^{2(+5e^-) = +10e^-}(aq) + CO_2(g)$$

$$MnO_4^-(aq) + \underbrace{C_2O_4^{2-}(aq) \longrightarrow Mn^{2+}(aq) + CO_2}_{5(-2e^-) = -10e^-}(g)$$

Notice that there are two carbon atoms, each of which lost one electron, making a total of two electrons. Place coefficients determined by the gain and loss of electrons:

$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) \rightarrow Mn^{2+}(aq) + CO_2(g)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the chromium and carbon atoms remain as above:

$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) \rightarrow$
$2Mn^{2+}(aq) + 10CO_2(g)$

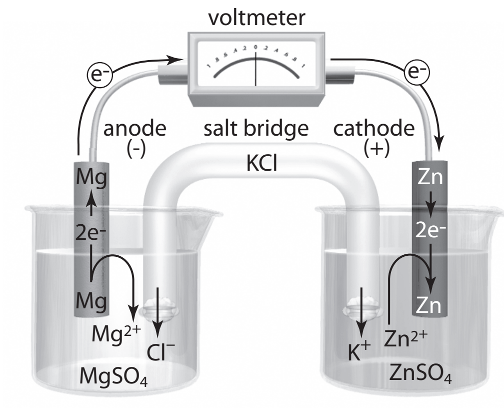
Add water to balance the oxygen atoms:

$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) \rightarrow$
$2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(\ell)$

Add hydrogen ions to balance the hydrogen atoms:

$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + 16H^+(aq) \rightarrow$
$2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(\ell)$

48.



49. Carbon monoxide is produced by heating carbon (in the form of coke) to 2000°C. This heated limestone and carbon monoxide that acts as the reducing agent for the ore, react with the iron(III) oxide in a series of processes that lead to the formation of liquid iron. The following are the reactions involved:

$3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$

$C(s) + CO_2(g) \rightarrow 2CO$

$FeO(s) + CO(g) \rightarrow Fe(\ell) + CO_2(g)$

50. After smelting, there are still some impurities in the iron. Purifying (or refining) this allows the iron (at this point called pig iron) into steel. Molten pig iron is poured into an upright vessel. Oxygen gas is pumped into the vessel over the iron. Lime, CaO(s), is then poured into the mixture. The impurities oxidize much more readily than does the iron. These oxidized compounds then react with the lime, according to the reactions shown here:

$SiO_2(s) + CaO(s) \rightarrow CaSiO_3(\ell)$

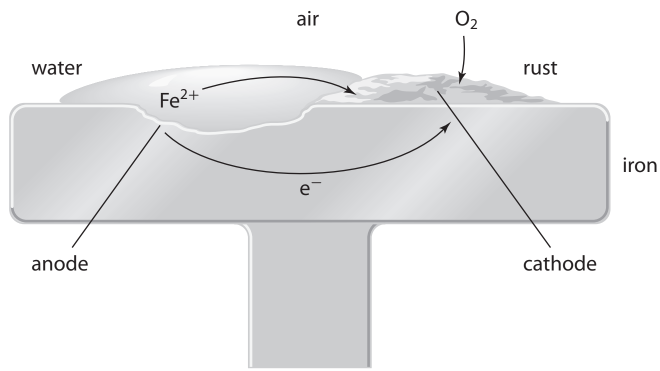
$P_4O_{10}(\ell) + 6CaO(s) \rightarrow 2Ca_3(PO_4)_2(\ell)$

The reactions produce a liquid called slag. The slag is less dense than the molten iron and therefore floats. When the reactions are complete, the vessel is tilted, and the slag poured off. The refined steel is then recovered.

51. a. Students should emphasize that oxidation occurs when an atom "loses" electrons and reduction occurs when an atom "gains" electrons, and that neither process occurs independently of the other. They may include sample half-reactions and redox equations showing oxidation numbers.

b. Galvanic cells convert chemical energy to electrical energy through spontaneous redox reactions. Electrolytic cells convert electrical energy to chemical energy by providing a source of energy to force a non-spontaneous redox reaction to occur. Students should include the idea that galvanic cells are essentially the reverse of electrolytic cells. Students should provide an example, including a diagram, of each type of cell.

52.



a. *Anode:* $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Cathode: $O_2(g) + 2H_2O(\ell) + 4e^- \rightarrow 4OH^-(aq)$

b. *Overall reaction:*

$2Fe(s) + O_2(g) + 2H_2O(\ell) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$

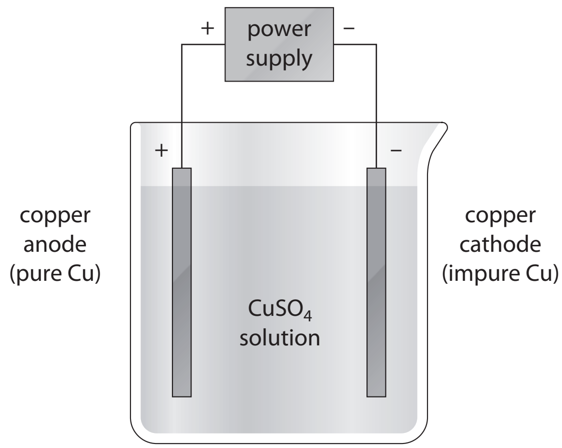
53. a. A disproportionation reaction occurs when one (or more) atom of an element is oxidized and another atom (or atoms) of the same element is reduced in the same reaction.

b. $3ClO^-(aq) \rightarrow 2Cl^-(aq) + ClO_3^-(aq)$

54. The ad should be written in plain language and emphasize the following:

- button batteries are easily swallowed
- if swallowed, button batteries can lodge in the esophagus
- body fluids will provide electrolytes necessary to activate the battery
- current will flow through the fluids, causing chemical reactions that produce strongly alkaline compounds, which will cause severe burns to the tissues within two hours of swallowing

55. a.



electrolytic refining of copper

Since the two electrodes are both made of copper, and neither water nor sulfate ions will create a potential difference in the cell, the cell potential is zero.

b. By adding an external source of electrons to create an anode and a cathode, the cell can be operated as an electrolytic cell. The copper metal anode will decrease in mass and the copper metal cathode will increase in mass.

56. Reports should include the type of fuel cells chosen for the station and why they were selected. There should be a little information about the mechanisms used by the fuel cells to produce electrical energy. The reports should also compare the amount of energy produced by the fuel cells with the energy produced by the solar arrays also used on the Space Station.

57. a. The solution would have a bluish colour to it, as copper(II) ions would now be in solution, and there would be spikes of silver (almost like a mossy covering) on the surface of the copper wire as the silver ions came out of solution as solid silver.

b. This is a redox reaction, as $Ag^{+}(aq)$ became $Ag(s)$ (a reduction) and $Cu(s)$ became $Cu^{2+}(aq)$ (an oxidation).

58. The graphic organizer should include the following steps:

Step 1: Write unbalanced half-reactions that show the formulas of the given reactant(s) and product(s).

Step 2: Balance any atoms other than oxygen and hydrogen first.

Step 3: Balance any oxygen atoms by adding water molecules.

Step 4: Balance any hydrogen atoms by adding hydrogen ions.

Step 5: Balance the charges by adding electrons.

59. Students should create a cyclic drawing of the flow of electrons as follows:

Start (anode) → wire → voltmeter → wire → cathode → ions in reduction half cell migrate to cathode where they accept the electrons and become metal atoms on the electrode

The electrons do not leave the cathode. The remainder of the current that completes the circuit is carried by ions in the solution.

60. A button battery is much smaller than an alkaline battery and thus can be used for things such as hearing aids, pacemakers, and smaller cameras and calculators. The two batteries have approx the same voltage of 1.5 V.

Alkaline battery:

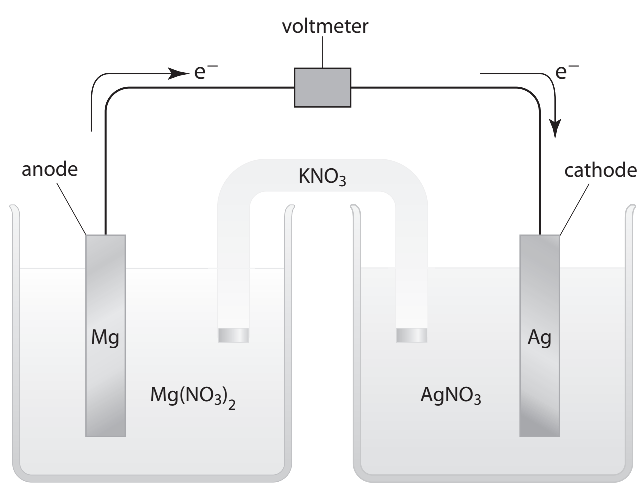
$Zn(s) + MnO_2(s) + H_2O(\ell) \rightarrow ZnO(s) + Mn(OH)_2(s)$

Button battery:

$Zn(s) + HgO(s) \rightarrow ZnO(s) + Hg(\ell)$

61. The Downs cell and the chlor-alkali process are similar in that they both produce chlorine gas through oxidation of chloride ions from sodium chloride. In the chlor-alkali process, a saturated aqueous solution of sodium chloride (brine) is electrolyzed. The products are chlorine gas, hydrogen gas, and sodium hydroxide. A Downs cell, in contrast, electrolyzes molten sodium chloride. The products are sodium metal and chlorine gas. To maintain the sodium chloride as a liquid, the reaction in a Downs cell must be carried out at a very high temperature (about 600°C). This is not necessary in the chlor-alkali process.

62. a to c



d. *Anode:* $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

Cathode: $Ag^{+}(aq) + e^- \rightarrow Ag(s)$

e. $Mg(s) + 2Ag^{+}(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$

f. $Mg^{2+} + 2e^- \rightarrow Mg(s)$ $E°_{anode} = -2.37$ V
$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E°_{cathode} = +0.80$ V
$E°_{cell} = E°_{cathode} - E°_{anode}$
$= +0.80 \text{ V} - (-2.37 \text{ V})$
$= 3.17$ V

g. $Mg(s) \mid Mg^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$

63. The internal make up of the cell must be such that the reversible process can occur safely and not overheat. The internal arrangement of materials in a rechargeable battery allow for the safe reversal of the process, which tends to be expensive in the manufacturing process. A disposable battery reduces cost by not taking the reversibility of the chemical process into account.

64. a. An electrolytic cell drives a non-spontaneous reaction, meaning that $E°$cell is negative. Since $E°_{cell} = E°_{cathode} - E°_{anode}$ and $E°_{anode}$ is 0 V, $E°_{cathode}$ must be negative.

b. Since $E°_{cell} = E°_{cathode} - E°_{anode}$ and $E°_{cathode}$ is 0 V, $E°_{anode}$ must be positive for $E°_{cell}$ to be negative.

65. For reactions that use only molecular compounds, the oxidation number method is easier.

66. All reports should include the concepts of cost for the lithium ion battery packs and the limited range of electric cars as well as their limited speed and pick up power in some situations.

67. The mass increase and the mass decrease are related through moles and, because the mass of a mole of one element is different from the mass of a mole of another element, there is no correlation between the mass increase or one electrode and the mass decrease in the other electrode.

68. The acidic solution of the tomatoes would cause the iron in the steel to rust. As a result, as the can ages, more iron would continue to move into solution, while weakening the can itself until the can fails to hold the liquid.

69. a. $6H_2O(\ell) \rightarrow$
$O_2(g) + 2H_2(g) + 4OH^-(aq) + 4H^+(aq)$

b. $E°_{cell} = E°_{cathode} - E°_{anode}$
$= -0.42 \text{ V} - 0.82 \text{ V}$
$= -1.24$ V

c. It is non-spontaneous.

d. The values do not represent standard conditions. It is not possible to create a 1.0 mol/L "solution" of water because pure water is about 55 mol/L.

70. Answers will depend on the biological process selected, but in all cases, the student must include a brief description of the process as well as the redox chemistry involved in the process.

Answers to Unit 5 Self-Assessment Questions (Student textbook pages 692-3)

1. e
2. b
3. c
4. e
5. a
6. d
7. a
8. d
9. d
10. c

11. Oxidation was originally defined as the reaction of an atom or compound with oxygen. Since then, oxidation has been redefined as a reaction in which electrons are lost.

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

Ionic equation:

$4Na(s) + O_2(g) \rightarrow 4Na^+(s) + 2O_2^-(s)$

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Ionic equation:

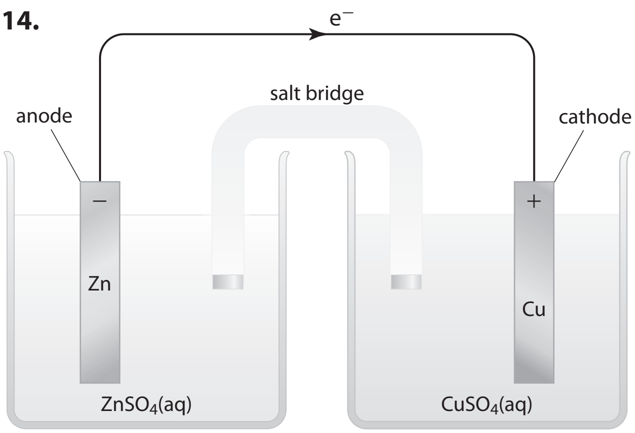
$2Na(s) + Cl_2(g) \rightarrow 2Na^+(s) + 2Cl^-(s)$

The modern definition of oxidation describes chemical reactions that resemble the reaction with oxygen, but with compounds other than oxygen. In both cases, electrons are lost by the sodium atoms.

12. a. N = +3
b. P = +5
c. Si = +4
d. N = −3, H = +1, S = +6, O = −2

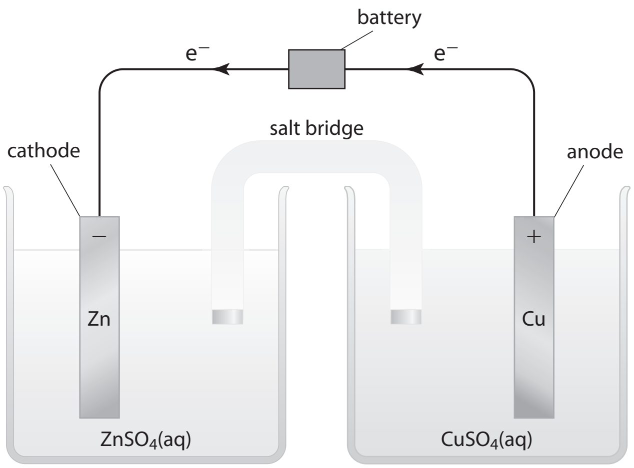
13. In this reaction, aluminum goes from an oxidation number of 0 to +3, indicating that is has been oxidized, while the oxygen goes from 0 to −2, indicating that it has been reduced.

14.



voltaic cell

b.



electrolytic cell

15. With moisture and salts within the ground, the iron in the steel pipe can corrode through oxidation. This would result in weakened areas along the pipeline that could eventually cause a leak. The zinc wire running alongside the steel pipe will act as a sacrificial anode and the zinc (a stronger reducing agent than the iron) will go through oxidation instead of the pipe. The zinc can periodically be replaced as it is used in the process of protecting the pipeline from corrosion.

16. a. $Hg_2^{\,2+}(aq) + 2e^- \rightarrow 2Hg$

b. $4H^+(aq) + TiO_2(s) + 2e^- \rightarrow Ti^{2+}(aq) + 2H_2O(\ell)$

c. $18OH^-(aq) + I_2(s) \rightarrow$
$2H_3IO_6^{\,3-}(aq) + 6H_2O(\ell) + 12e^-$

17. No. An oxidation number of −2 for oxygen in a water molecule means that the oxygen atom attracts electrons more strongly than the hydrogen atom. That is, oxygen has a higher electronegativity than hydrogen. Thus, the electrons that are shared by the oxygen atom and the hydrogen atoms in a water molecule are all considered to "belong" to the oxygen atom. Thus, when a water molecule forms, the oxygen atom is considered to have gained two electrons and each hydrogen atom is considered to have lost an electron.

18. b. aqueous nickel sulfate and metallic zinc

19. The graphic organizer should include the following steps:

Step 1: Write unbalanced half-reactions that show the formulas of the given reactant(s) and product(s).

Step 2: Balance any atoms other than oxygen and hydrogen first.

Step 3: Balance any oxygen atoms by adding water molecules.

Step 4: Balance any hydrogen atoms by adding hydrogen ions.

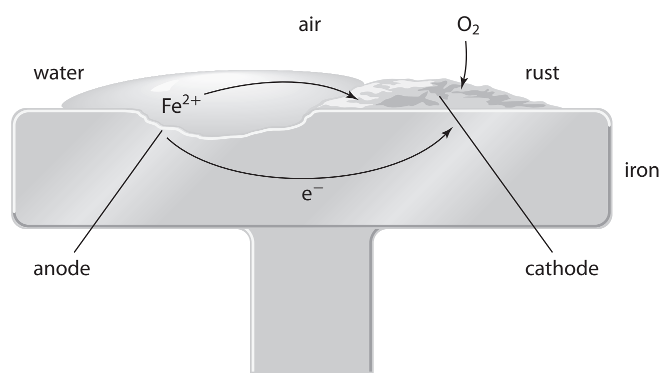
Step 5: Adjust for basic conditions by adding to both sides the same number of hydroxide ions as the number of hydrogen ions already present.

Step 6: Simplify the equation by combining hydrogen ions and hydroxide ions that appear on the same side of the equation into water molecules.

Step 7: Cancel any water molecules present on both sides of the equation.

Step 8: Balance the charges by adding electrons.

20.



21. a. Assign oxidation numbers:

$$\underset{+3}{Cr}^{3+}(aq) + \underset{+5\ -2}{ClO_3^-}(aq) \longrightarrow \underset{+6\ -2}{CrO_4^{2-}}(aq) + \underset{-1}{Cl^-}(aq)$$

Track electrons:

$$\overbrace{Cr^{3+}(aq) + ClO_3^-(aq) \longrightarrow CrO_4^{2-}}^{2(-3e^-) = -6e^-}(aq) + Cl^-(aq)$$

$$Cr^{3+}(aq) + \underbrace{ClO_3^-(aq) \longrightarrow CrO_4^{2-}(aq) + Cl^-}_{1(+6e^-) = +6e^-}(aq)$$

Place coefficients determined by the gain and loss of electrons:

$2Cr^{3+}(aq) + ClO_3^-(aq) \rightarrow CrO_4^{2-}(aq) + Cl^-(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the chromium and chlorine atoms remain as above:

$2Cr^{3+}(aq) + ClO_3^-(aq) \rightarrow 2CrO_4^{2-}(aq) + Cl^-(aq)$

Add water to balance the oxygen atoms:

$2Cr^{3+}(aq) + ClO_3^-(aq) + 5H_2O(\ell) \rightarrow$
$2CrO_4^{2-}(aq) + Cl^-(aq)$

Add hydrogen ions to balance the hydrogen atoms:

$2Cr^{3+}(aq) + ClO_3^-(aq) + 5H_2O(\ell) \rightarrow$
$2CrO_4^{2-}(aq) + Cl^-(aq) + 10H^+(aq)$

b. Assign oxidation numbers:

$$\underset{+4\ -2}{ClO_2}(g) + \underset{+3\ -2}{SbO_2^-}(aq) \longrightarrow \underset{+6\ -2}{ClO_2^-}(aq) + \underset{+5\ \ -2\ +1}{Sb(OH)_6^-}(aq)$$

Track electrons:

$2(+1e^-) = +2e^-$

$$ClO_2(g) + SbO_2^-(aq) \longrightarrow ClO_2^-(aq) + Sb(OH)_6^-(aq)$$

$1(-2e^-) = -2e^-$

Place coefficients determined by the gain and loss of electrons:

$2ClO_2(g) + SbO_2^-(aq) \rightarrow$
$ClO_2^-(aq) + Sb(OH)_6^-(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the chlorine and antimony atoms remain as above:

$2ClO_2(g) + SbO_2^-(aq) \rightarrow$
$2ClO_2^-(aq) + Sb(OH)_6^-(aq)$

Add water to balance the oxygen atoms:

$2ClO_2(g) + SbO_2^-(aq) + 4H_2O(\ell) \rightarrow$
$2ClO_2^-(aq) + Sb(OH)_6^-(aq)$

Add hydrogen ions to balance the hydrogen atoms:

$2ClO_2(g) + SbO_2^-(aq) + 4H_2O(\ell) \rightarrow$
$2ClO_2^-(aq) + Sb(OH)_6^-(aq) + 2H^+(aq)$

To both sides, add the number of hydroxide ions equal to the number of hydrogen ions.

$2ClO_2(g) + SbO_2^-(aq) + 4H_2O(\ell) + 2OH^-(aq) \rightarrow$
$2ClO_2^-(aq) + Sb(OH)_6^-(aq) + 2H^+(aq)$
$+ 2OH^-(aq)$

Combine H^+ and OH^- to make water, where possible.

$2ClO_2(g) + SbO_2^-(aq) + 4H_2O(\ell) + 2OH^-(aq) \rightarrow$
$2ClO_2^-(aq) + Sb(OH)_6^-(aq) + 2H_2O(\ell)$

Cancel water molecules that are on both sides of the equation and check to ensure that atoms of all elements are balanced.

$2ClO_2(g) + SbO_2^-(aq) + 2H_2O(\ell) + 2OH^-(aq) \rightarrow$
$2ClO_2^-(aq) + Sb(OH)_6^-(aq)$

c. Assign oxidation numbers:

$$\underset{+1\ +5\ -2}{HNO_3}(aq) + \underset{+1\ +3\ -2}{H_3AsO_3}(aq) \longrightarrow \underset{+2\ -2}{NO}(g) + \underset{+1\ +5\ -2}{H_3AsO_4}(aq)$$

Track electrons:

$2(+3e^-) = +6e^-$

$$HNO_3(aq) + H_3AsO_3(aq) \longrightarrow NO(g) + H_3AsO_4(aq)$$

$3(-2e^-) = -6e^-$

Place coefficients determined by the gain and loss of electrons:

$2HNO_3(aq) + 3H_3AsO_3(aq) \rightarrow$
$NO(g) + H_3AsO_4(aq)$

Balance everything except oxygen and hydrogen while ensuring that the ratios of the nitrogen and arsenic atoms remain as above:

$2HNO_3(aq) + 3H_3AsO_3(aq) \rightarrow$
$2NO(g) + 3H_3AsO_4(aq)$

Add water to balance the oxygen atoms:

$2HNO_3(aq) + 3H_3AsO_3(aq) \rightarrow$
$2NO(g) + 3H_3AsO_4(aq) + H_2O(\ell)$

This obviously occurs in an acidic solution because of the presence of nitric acid in the reactants. However, the hydrogen ions are combined with the nitrate ion so just check to see that the hydrogen atoms are balanced.

There are two hydrogen atoms on each side so the equation is balanced.

22. Due to the high demand for sodium hydroxide and chlorine gas for many industrial processes and in the manufacturing of many materials used every day, the chlor-alkali process is the important industrial process where an aqueous solution of sodium hydroxide is electrolyzed to produce chlorine gas, and sodium hydroxide. Hydrogen gas is also a product of the process.

23. $Sn(s) \mid Sn^{2+}(aq) \parallel Fe^{3+}(aq), Fe^{2+}(aq) \mid Pt(s)$

24. The mass of an anode always decreases when the material is involved in the oxidation half-reaction. This is due to the fact that at the anode, the metal atoms are oxidized and becomes ions that go into solution. Likewise, at the cathode, ions in the solution are reduced (accept electrons) and become metal atoms that are deposited on the cathode. As a result, the cathode mass increases. Therefore, the anode in the process was electrode 2 and electrode 1 was the cathode.

25. a. *Reactants*: sodium chloride (and calcium chloride to simply lower the melting point, so it is not really considered a reactant).

Products: chlorine gas, liquid sodium

b. *Anode*: $2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

Cathode: $Na^+(\ell) + e^- \rightarrow Na(\ell)$