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

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

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

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

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

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