CHAPTER 7

Reactions in Aqueous Solution

CHAPTER ANSWERS

1. Water is the most universal of all liquids. Water has a relatively large heat capacity and a relatively large liquid range, which means it can absorb the heat liberated by many reactions while still remaining in the liquid state. Water is very polar and dissolves well both ionic solutes and solutes with which it can hydrogen bond (this is especially important to the biochemical reactions of the living cell).

2. Driving forces are types of *changes* in a system that pull a reaction in the *direction of product formation*; driving forces discussed in Chapter 7 are formation of a *solid*, formation of *water*, formation of a *gas*, and transfer of electrons.

3. A *precipitation reaction* is one in which a *solid* forms (most typically when solutions of two ionic solutes are mixed).

4. The net charge of a precipitate must be *zero*. The total number of positive charges equals the total number of negative charges.

5. When an ionic solute such as NaCl (sodium chloride) is dissolved in water, the resulting solution consists of separate, individual, discrete hydrated sodium ions (Na^+) and separate, individual, discrete hydrated chloride ions (Cl^-). There are no identifiable NaCl units in such a solution, and the positive and negative ions behave independently of one another.

6. ions

7. A substance is said to be a strong electrolyte if *each* unit of the substance produces separate, distinct ions when the substance dissolves in water. NaCl and KNO_3 are both strong electrolytes.

8. Chemists know when a solution contains independent ions because such a solution will readily allow an electrical current to pass through it. The simplest experiment that demonstrates this uses the sort of light–bulb conductivity apparatus described in the text; if the light bulb glows strongly, then the solution must contain a lot of ions that conduct the electricity well.

9. $NaNO_3$ must be soluble in water.

10. Answer depends on student choices.

11.
    a. soluble (Rule 1: Most nitrate salts are soluble.)
    b. soluble (Rule 3: Most chloride salts are soluble.)
    c. soluble (Rule 4: Most sulfate salts are soluble.)
    d. insoluble (Rule 5: Most hydroxide compounds are insoluble.)
    e. insoluble (Rule 6: Most sulfide salts are insoluble.)
    f. insoluble (Rule 5: Most hydroxide compounds are insoluble.)
    g. insoluble (Rule 6: Most phosphate salts are insoluble.)

12.

a. soluble; Rule 3

b. Rule 6: Most sulfide salts are only slightly soluble.

c. Rule 5: Most hydroxides are only slightly soluble.

d. soluble; Rule 2

e. soluble; Rule 4

f. Rule 6: Most sulfide salts are only slightly soluble.

g. soluble; Rule 2

h. Rule 6: Most carbonate salts are only slightly soluble

13.

a. Rule 6: Most sulfide salts are only slightly soluble.

b. Rule 5: Most hydroxide compounds are only slightly soluble.

c. Rule 6: Most phosphate salts are only slightly soluble.

d. Rule 3, exception to the rule for chlorides

14.

a. Rule 5: Most hydroxides are only slightly soluble.

b. Rule 6: Most phosphate salts are only slightly soluble.

c. Rule 6: Most carbonate salts are only slightly soluble.

d. Rule 6: Most sulfide salts are only slightly soluble.

15.

a. $FePO_4$; Rule 6: Most phosphate salts are only slightly soluble.

b. $BaSO_4$; Rule 4, exception to the rule for sulfates

c. No precipitate is likely; Rules 2, 3, and 4.

d. $PbCl_2$; Rule 3: $PbCl_2$ is a listed exception.

e. No precipitate is likely; Rules 1, 2, and 3.

f. CuS; Rule 6: Most sulfide salts are only slightly soluble.

16.

a. $MnCO_3$; Rule 6: Most carbonates are only slightly soluble.

b. $CaSO_4$; Rule 4, exception for sulfates

c. Hg_2Cl_2; Rule 3, exception for chlorides

d. soluble

e. $Ni(OH)_2$; Rule 5: Most hydroxides are only slightly soluble.

f. $BaSO_4$; Rule 4, exception for sulfates

17. The precipitates are marked in boldface type.

   a. No precipitate; both $(NH_4)_2SO_4$ and HCl are soluble.
   $NH_4Cl(aq) + H_2SO_4(aq) \rightarrow$ no precipitate

   b. Rule 6: Most carbonate salts are only slightly soluble.
   $2K_2CO_3(aq) + SnCl_4(aq) \rightarrow \mathbf{Sn(CO_3)_2}(s) + 4KCl(aq)$

   c. Rule 3; exception to rule for chlorides
   $2NH_4Cl(aq) + Pb(NO_3)_2(aq) \rightarrow \mathbf{PbCl_2}(s) + 2NH_4NO_3(aq)$

   d. Rule 5: Most hydroxide compounds are only slightly soluble.
   $CuSO_4(aq) + 2KOH(aq) \rightarrow \mathbf{Cu(OH)_2}(s) + K_2SO_4(aq)$

   e. Rule 6: Most phosphate salts are only slightly soluble.
   $Na_3PO_4(aq) + CrCl_3(aq) \rightarrow \mathbf{CrPO_4}(s) + 3NaCl(s)$

   f. Rule 6: Most sulfide salts are only slightly soluble.
   $3(NH_4)_2S(aq) + 2FeCl_3(aq) \rightarrow \mathbf{Fe_2S_3}(s) + 6NH_4Cl(aq)$

18. The precipitates are marked in boldface type.

   a. Rule 6: Most sulfide salts are insoluble. $Na_2S(aq) + CuCl_2(aq)$ $\mathbf{CuS}(s) + 2NaCl(aq)$

   b. Rule 6: Most phosphate salts are insoluble. $K_3PO_4(aq) + AlCl_3(aq)$ $3KCl(aq) + \mathbf{AlPO_4}(s)$

   c. Rule 4: barium sulfate is a listed exception. $H_2SO_4(aq) + BaCl_2(aq)$ $\mathbf{BaSO_4}(s) + 2HCl(aq)$

   d. Rule 5: Most hydroxide compounds are insoluble. $3NaOH(aq) + FeCl_3(aq)$ $3NaCl(aq) + \mathbf{Fe(OH)_3}(s)$

   e. Rule 3: a listed exception for chlorides $2NaCl(aq) + Hg_2(NO_3)_2(aq)$ $2NaNO_3(aq) + \mathbf{Hg_2Cl_2}(s)$

   f. Rule 6: Most carbonate salts are insoluble. $3K_2CO_3(aq) + 2Cr(C_2H_3O_2)_3(aq)$ $6KC_2H_3O_2(aq) + \mathbf{Cr_2(CO_3)_3}(s)$

19. Hint: When balancing equations involving polyatomic ions, especially in precipitation reactions, balance the polyatomic ions as a *unit*, not in terms of the atoms the polyatomic ions contain (e.g., treat nitrate ion, NO_3^-, as a single entity, not as one nitrogen and three oxygen atoms). When finished balancing, however, be sure to count the individual number of atoms of each type on each side of the equation.

   a. $Na_2SO_4(aq) + CaCl_2(aq) \rightarrow CaSO_4(s) + NaCl(aq)$
   Balance sodium: $Na_2SO_4(aq) + CaCl_2(aq) \rightarrow CaSO_4(s) + 2NaCl(aq)$
   Balanced equation: $Na_2SO_4(aq) + CaCl_2(aq) \rightarrow CaSO_4(s) + 2NaCl(aq)$

   b. $Co(C_2H_3O_2)_2(aq) + Na_2S(aq) \rightarrow CoS(s) + NaC_2H_3O_2(aq)$
   Balance acetate: $Co(C_2H_3O_2)_2(aq) + Na_2S(aq) \rightarrow CoS(s) + 2NaC_2H_3O_2(aq)$
   Balanced equation: $Co(C_2H_3O_2)_2(aq) + Na_2S(aq) \rightarrow CoS(s) + 2NaC_2H_3O_2(aq)$

c. $KOH(aq) + NiCl_2(aq) \rightarrow Ni(OH)_2(s) + KCl(aq)$

Balance hydroxide: $2KOH(aq) + NiCl_2(aq) \rightarrow Ni(OH)_2(s) + KCl(aq)$

Balance potassium: $2KOH(aq) + NiCl_2(aq) \rightarrow Ni(OH)_2(s) + 2KCl(aq)$

Balanced equation: $2KOH(aq) + NiCl_2(aq) \rightarrow Ni(OH)_2(s) + 2KCl(aq)$

20. Hint: When balancing equations involving polyatomic ions, especially in precipitation reactions, balance the polyatomic ions as a *unit*, not in terms of the atoms the polyatomic ions contain (e.g., treat nitrate ion, NO_3^-, as a single entity, not as one nitrogen and three oxygen atoms). When finished balancing, however, be sure to count the individual number of atoms of each type on each side of the equation.

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

balance chlorine: $CaCl_2(aq) + AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$

balance silver: $CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$

balanced equation: $CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$

b. $AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + KNO_3(aq)$

balance silver: $2AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + KNO_3(aq)$

balance nitrate ion: $2AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2KNO_3(aq)$

balanced equation: $2AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2KNO_3(aq)$

c. $BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + KCl(aq)$

balance potassium: $BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KCl(aq)$

balanced equation: $BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KCl(aq)$

21. The products are determined by having the ions "switch partners." For example, for a general reaction AB + CD →, the possible products are AD and CB if the ions switch partners. If either AD or CB is insoluble, then a precipitation reaction has occurred. In the following reaction, the formula of the precipitate is given in boldface type.

a. $Ba(NO_3)_2(aq) + (NH_4)_2SO_4(aq) \rightarrow \mathbf{BaSO_4}(s) + 2NH_4NO_3(aq)$

Rule 4: $BaSO_4$ is a listed exception.

b. $CoCl_3(aq) + 3NaOH \rightarrow \mathbf{Co(OH)_3}(s) + 3NaCl(aq)$

Rule 5: Most hydroxide compounds are only slightly soluble.

c. $2FeCl_3(aq) + 3(NH_4)_2S(aq) \rightarrow \mathbf{Fe_2S_3}(s) + 6NH_4Cl(aq)$

Rule 6: Most sulfide salts are only slightly soluble.

22. The products are determined by having the ions "switch partners." For example, for a general reaction AB + CD →, the possible products are AD and CB if the ions switch partners. If either AD or CB is insoluble, then a precipitation reaction has occurred. In the following reaction, the formula of the precipitate is given in boldface type.

a. $CaCl_2(aq) + 2AgC_2H_3O_2(aq) \rightarrow \mathbf{2AgCl}(s) + Ca(C_2H_3O_2)_2(aq)$

Rule 3; exception for chloride

b. $Ba(NO_3)_2(aq) + 2NH_4OH(aq) \rightarrow$ **$Ba(OH)_2(s)$** $+ 2NH_4NO_3(aq)$

Rule 5: Most hydroxides are only slightly soluble.

c. $NiCl_2(aq) + Na_2CO_3(aq) \rightarrow$ **$NiCO_3(s)$** $+ 2NaCl(aq)$

Rule 6: Most carbonates are only slightly soluble.

23. The *net ionic equation* for a reaction in solution indicates only those components that are directly involved in the reaction. Other ions that may be present to balance charge, but which do not actively participate in the reaction are called *spectator ions* and are not indicated when writing the chemical equation for the reaction.

24. Spectator ions are ions that *remain in solution* during a precipitation/double displacement reaction. For example in the reaction

$BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KCl(aq)$

the K^+ and Cl^- ions are the spectator ions.

25. The net ionic equation for a reaction indicates *only those ions that form the precipitate* and does not show the spectator ions present in the solutions mixed. The identity of the precipitate is determined from the Solubility Rules (Table 7.1).

a. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Rule 3: AgCl is listed as an insoluble exception.

b. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Rule 4: $BaSO_4$ is listed as an insoluble exception.

c. $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

Rule 6: Most phosphate salts are only slightly soluble.

d. both KF and H_2SO_4 are soluble; no precipitate

e. $Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$

Rule 4: $CaSO_4$ is listed as an insoluble exception.

f. $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

Rule 3: $PbCl_2$ is listed as an insoluble exception.

26. The net ionic equation for a reaction indicates *only those ions that go to form the precipitate* and does not show the spectator ions present in the solutes mixed. The identity of the precipitate is determined from the Solubility Rules (Table 7.1).

a. $Ca(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(s) + 2HNO_3(aq)$

$Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$

b. $Ni(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Ni(OH)_2(s) + 2NaNO_3(aq)$

$Ni^{2+}(aq) + 2OH^-(aq) \rightarrow Ni(OH)_2(s)$

c. $3(NH_4)_2S(aq) + 2FeCl_3(aq) \rightarrow Fe_2S_3(s) + 6NH_4Cl(aq)$

$2Fe^{3+}(aq) + 3S^{2-}(aq) \rightarrow Fe_2S_3(s)$

27. $Cu^{2+}(aq) + CrO_4^{2-}(aq) \rightarrow CuCrO_4(s)$
$Co^{3+}(aq) + CrO_4^{2-}(aq) \rightarrow Co_2(CrO_4)_3(s)$
$Ba^{2+}(aq) + CrO_4^{2-}(aq) \rightarrow BaCrO_4(s)$
$Fe^{3+}(aq) + CrO_4^{2-}(aq) \rightarrow Fe_2(CrO_4)_3(s)$

28. $Ag^{+}(aq) + Cl^{-}(aq) \rightarrow AgCl(s)$
$Pb^{2+}(aq) + 2Cl^{-}(aq) \rightarrow PbCl_2(s)$
$Hg_2^{2+}(aq) + 2Cl^{-}(aq) \rightarrow Hg_2Cl_2(s)$

29. $Ca^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow CaC_2O_4(s)$

30. $Co^{2+}(aq) + S^{2-}(aq) \rightarrow CoS(s)$
$2Co^{3+}(aq) + 3S^{2-}(aq) \rightarrow Co_2S_3(s)$
$Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
$2Fe^{3+}(aq) + 3S^{2-}(aq) \rightarrow Fe_2S_3(s)$

31. Strong acids ionize completely in water. The strong acids are also strong electrolytes.

32. Strong bases fully produce hydroxide ions when dissolved in water. The strong bases are also strong electrolytes.

33. $H^{+}(aq) + OH^{-}(aq) \rightarrow H_2O$; formation of a water molecule

34. acids: HCl, H_2SO_4, HNO_3, $HClO_4$, HBr
bases: NaOH, KOH, RbOH, CsOH

35. 1000; 1000

36. A salt is the ionic product remaining in solution when an acid neutralizes a base. For example, in the reaction $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$, sodium chloride is the salt produced by the neutralization reaction.

37. Your textbook mentions four strong acids. You had to give only three of the following equations.
$HCl(aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$
$HNO_3(aq) \rightarrow H^{+}(aq) + NO_3^{-}(aq)$
$H_2SO_4(aq) \rightarrow H^{+}(aq) + HSO_4^{-}(aq)$
$HClO_4(aq) \rightarrow H^{+}(aq) + ClO_4^{-}(aq)$

38. $RbOH(s) \rightarrow Rb^{+}(aq) + OH^{-}(aq)$
$CsOH(s) \rightarrow Cs^{+}(aq) + OH^{-}(aq)$

39. The formulas of the salts are marked in boldface type. Remember that in an acid/base reaction in aqueous solution, *water* is always one of the products; keeping this in mind makes predicting the formula of the *salt* produced easy to do.

a. $HCl(aq) + KOH(aq) \rightarrow H_2O(l) + \mathbf{KCl}(aq)$

b. $RbOH(aq) + HNO_3(aq) \rightarrow H_2O(l) + \mathbf{RbNO_3}(aq)$

c. $HClO_4(aq) + NaOH(aq) \rightarrow H_2O(l) + \mathbf{NaClO_4}(aq)$

d. $HBr(aq) + CsOH(aq) \rightarrow H_2O(l) + CsBr(aq)$

40. In general, the salt formed in an aqueous acid–base reaction consists of the *positive ion of the base* involved in the reaction combined with the *negative ion of the acid*. The hydrogen ion of the strong acid combines with the hydroxide ion of the strong base to produce water, which is the other product of the acid–base reactions.

a. $KOH(aq) + HCl(aq) \rightarrow H_2O(l) + KCl(aq)$

b. $NaOH(aq) + HClO_4(aq) \rightarrow H_2O(l) + NaClO_4(aq)$

c. $CsOH(aq) + HNO_3(aq) \rightarrow H_2O(l) + CsNO_3(aq)$

d. $2KOH(aq) + H_2SO_4(aq) \rightarrow 2H_2O(l) + K_2SO_4(aq)$

41. An oxidation–reduction reaction is one in which one species loses electrons (oxidation) and another species gains electrons (reduction). Electrons are transferred from the species being oxidized to the species being reduced.

42. Answer depends on student choice of example: $Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$ is an example.

43. A driving force, in general, is an event that tends to help convert the reactants of a process into the products. Some elements (metals) tend to lose electrons whereas other elements (nonmetals) tend to gain electrons. A *transfer* of electrons from atoms of a metal to atoms of a nonmetal would be favorable and would result in a chemical reaction. A simple example of such a process is the reaction of sodium with chlorine; sodium atoms tend to each lose one electron (to form Na^+) whereas chlorine atoms tend to each gain one electron (to form Cl^-). The reaction of sodium metal with chlorine gas represents a transfer of electrons from sodium atoms to chlorine atoms to form sodium chloride.

44. The metallic element *loses* electrons, and the nonmetallic element *gains* electrons.

45. Each calcium atom would lose two electrons. Each fluorine atom would gain one electron (so the F_2 molecule would gain two electrons). One calcium atom would be required to react with one fluorine, F_2, molecule. Calcium ions are charged 2+; fluoride ions are charged 1–.

46. Each magnesium atom would lose two electrons. Each oxygen atom would gain two electrons (so the O_2 molecule would gain four electrons). Two magnesium atoms would be required to react with each oxygen, O_2, molecule. Magnesium ions are charged 2+; oxide ions are charged 2–.

47. $MgCl_2$ is made up of Mg^{2+} ions and Cl^- ions. Magnesium atoms each lose twoelectrons to become Mg^{2+} ions. Chlorine atoms each gain one electron to become Cl^- ions (so each Cl_2 molecule gains two electrons to become two Cl^- ions).

48. Each potassium atom loses one electron. The sulfur atom gains two electrons. So, two potassium atoms are required to react with one sulfur atom.

$2 \times (K \rightarrow K^+ + e^-)$

$S + 2e^- \rightarrow S^{2-}$

49.

a. $Na + S \rightarrow Na_2S$

Balance sodium: $2Na + S \rightarrow Na_2S$

Balanced equation: $2Na(s) + S(s) \rightarrow Na_2S(s)$

sodium is oxidized, sulfur is reduced

b. $Mg + O_2 \rightarrow MgO$

Balance oxygen: $Mg + O_2 \rightarrow 2MgO$

Balance magnesium: $2Mg + O_2 \rightarrow 2MgO$

Balanced equation: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

magnesium is oxidized, oxygen is reduced

c. Equation is already balanced.

$Ca(s) + F_2(s) \rightarrow CaF_2(s)$

calcium is oxidized, fluorine is reduced

d. $Fe + Cl_2 \rightarrow FeCl_3$

Balance chlorine: $Fe + 3Cl_2 \rightarrow 2FeCl_3$

Balance iron: $2Fe + 3Cl_2 \rightarrow 2FeCl_3$

Balanced equation: $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$

iron is oxidized, chlorine is reduced

50.

a. $P_4(s) + O_2(g) \rightarrow P_4O_{10}(s)$

balance oxygen: $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$

balanced equation: $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$

b. $MgO(s) + C(s) \rightarrow Mg(s) + CO(g)$

This equation is already balanced.

c. $Sr(s) + H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$

balance oxygen: $Sr(s) + 2H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$

balanced equation: $Sr(s) + 2H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$

d. $Co(s) + HCl(aq) \rightarrow CoCl_2(aq) + H_2(g)$

balance hydrogen: $Co(s) + 2HCl(aq) \rightarrow CoCl_2(aq) + H_2(g)$

balanced equation: $Co(s) + 2HCl(aq) \rightarrow CoCl_2(aq) + H_2(g)$

51. In a double-displacement reaction, two ionic solutes "switch partners" with the positive ion from one combining with the negative ion from the other to form the precipitate. For example, in the reaction $AgNO_3(aq) + HCl(aq) \rightarrow AgCl(s) + HNO_3(aq)$, silver ion from one solute combines with chloride ion from the other solute to form the precipitate. In a single displacement reaction, one element replaces another from its compound; in other words, a single displacement reaction is typically an oxidation reduction reaction. Also for example, in the reaction $Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$, zinc in the elemental form replaces copper in the copper compound, producing copper in the elemental form and a zinc compound.

52. Examples of formation of water:

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow 2H_2O(l) + K_2SO_4(aq)$$

Examples of formation of a gaseous product:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

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

53. For each reaction, the type of reaction is first identified, followed by some of the reasoning that leads to this choice (there may be more than one way in which you can recognize a particular type of reaction).

a. precipitation (From Table 7.1, $BaSO_4$ is insoluble).

b. oxidation–reduction (Zn changes from the elemental to the combined state; hydrogen changes from the combined to the elemental state).

c. precipitation (From Table 7.1, AgCl is insoluble.)

d. acid–base (HCl is an acid; KOH is a base; water and a salt are produced.)

e. oxidation–reduction (Cu changes from the combined to the elemental state; Zn changes from the elemental to the combined state.)

f. acid–base (The $H_2PO_4^-$ ion behaves as an acid; NaOH behaves as a base; a salt and water are produced.)

g. precipitation (From Table 7.1, $CaSO_4$ is insoluble); acid–base [$Ca(OH)_2$ is a base; H_2SO_4 is an acid; a salt and water are produced.]

h. oxidation–reduction (Mg changes from the elemental to the combined state; Zn changes from the combined to the elemental state.)

i. precipitation (From Table 7.1, $BaSO_4$ is insoluble.)

54. For each reaction, the type of reaction is first identified, followed by some of the reasoning that leads to this choice (There may be more than one way in which you can recognize a particular type of reaction.).

a. oxidation–reduction (Oxygen changes from the combined state to the elemental state.)

b. oxidation–reduction (Zinc changes from the elemental to the combined state; hydrogen changes from the combined to the elemental state.)

c. acid–base (H_2SO_4 is a strong acid, and NaOH is a strong base; water and a salt are formed.)

d. acid–base, precipitation [H_2SO_4 is a strong acid, and $Ba(OH)_2$ is a base; water and a salt are formed; an insoluble product forms.]

e. precipitation (From the Solubility Rules of Table 7.1, AgCl is only slightly soluble.)

f. precipitation [From the Solubility Rules of Table 7.1, $Cu(OH)_2$ is only slightly soluble.]

g. oxidation–reduction (Chlorine and fluorine change from the elemental to the combined state.)

h. oxidation–reduction (Oxygen changes from the elemental to the combined state.)

i. acid–base (HNO_3 is a strong acid and $Ca(OH)_2$ is a strong base; a salt and water are formed.)

55. A combustion reaction is typically a reaction in which an element or compound reacts with oxygen so quickly and with so much release of energy that a flame results. In addition to the carbon dioxide and water chemical products, combustion reactions are a major source of heat energy.

56. oxidation–reduction

57. A synthesis reaction represents the production of a given compound from simpler substances (either elements or simpler compounds). For example,

$O_2(g) + 2F_2(g) \rightarrow 2OF_2(g)$

represents a simple synthesis reaction. Synthesis reactions may often (but not necessarily always) also be classified in other ways. For example, the reaction

$C(s) + O_2(g) \rightarrow CO_2(g)$

could also be classified as an oxidation–reduction reaction or as a combustion reaction (a special subclassification of oxidation–reduction reaction that produces a flame). As another example, the reaction

$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$

is a synthesis reaction that is also an oxidation–reduction reaction.

58. A decomposition reaction is one in which a given compound is broken down into simpler compounds or constituent elements. The reactions

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

$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

both represent decomposition reactions. Such reactions often (but not necessarily always) may be classified in other ways. For example, the reaction of $HgO(s)$ is also an oxidation–reduction reaction.

59. Compounds like those in this problem, which contain only carbon and hydrogen, are called *hydrocarbons*. When a hydrocarbon is reacted with oxygen (O_2), the hydrocarbon is almost always converted to carbon dioxide and water vapor. Because water molecules contain an odd number of oxygen atoms, and O_2 contains an even number of oxygen atoms, it is often difficult to balance such equations. For this reason, it is simpler to balance the equation using fractional coefficients if necessary and then multiply by a factor that will give whole number coefficients for the final balanced equation.

a. $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

Balance carbon: $C_2H_6 + O_2 \rightarrow 2CO_2 + H_2O$

Balance hydrogen: $C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$

Balance oxygen: $C_2H_6 + (7/2)O_2 \rightarrow 2CO_2 + 3H_2O$

Balanced equation: $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

b. $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

Balance carbon: $C_4H_{10} + O_2 \rightarrow 4CO_2 + H_2O$

Balance hydrogen: $C_4H_{10} + O_2 \rightarrow 4CO_2 + 5H_2O$

Balance oxygen: $C_4H_{10} + (13/2)O_2 \rightarrow 4CO_2 + 5H_2O$

Balanced equation: $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

c. $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$

Balance carbon: $C_6H_{14} + O_2 \rightarrow 6CO_2 + H_2O$

Balance hydrogen: $C_6H_{14} + O_2 \rightarrow 6CO_2 + 7H_2O$

Balance oxygen: $C_6H_{14} + (19/2)O_2 \rightarrow 6CO_2 + 7H_2O$

Balanced equation: $2C_6H_{14}(g) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

60. Compounds like those in this problem, which contain only carbon and hydrogen, are called ***hydrocarbons***. When a hydrocarbon is reacted with oxygen (O_2), the hydrocarbon is almost always converted to carbon dioxide and water vapor. Because water molecules contain an odd number of oxygen atoms, and O_2 contains an even number of oxygen atoms, it is often difficult to balance such equations. For this reason, it is simpler to balance the equation using fractional coefficients if necessary and then multiply by a factor that will give whole number coefficients for the final balanced equation.

a. $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

balance carbon: $C_3H_8(g) + O_2(g) \rightarrow 3CO_2(g) + H_2O(g)$

balance hydrogen: $C_3H_8(g) + O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

balance oxygen: $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

balanced equation: $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

b. $C_2H_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

balance carbon: $C_2H_4(g) + O_2(g) \rightarrow 2CO_2(g) + H_2O(g)$

balance hydrogen: $C_2H_4(g) + O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

balance oxygen: $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

balanced equation: $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

c. $C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

balance carbon: $C_4H_{10}(g) + O_2(g) \rightarrow 4CO_2(g) + H_2O(g)$

balance hydrogen: $C_4H_{10}(g) + O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

balance oxygen: $C_4H_{10}(g) + (13/2)O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

balanced equation: $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

61. Specific examples will depend on the students' input. A typical combustion reaction is represented by the reaction of methane (CH_4) with oxygen gas,

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

62. A reaction in which small molecules or atoms combine to make a larger molecule is called a *synthesis* reaction. An example would be the synthesis of sodium chloride from the elements,

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$.

A reaction in which a molecule is broken down into simpler molecules or atoms is called a *decomposition* reaction. An example would be the decomposition of sodium hydrogen carbonate when heated,

$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$.

Specific examples will depend on the students' input.

63.

a. $Ni(s) + 4CO(g) \rightarrow Ni(CO)_4(g)$

b. $2Al(s) + 3S(s) \rightarrow Al_2S_3(s)$

c. $Na_2SO_3(aq) + S(s) \rightarrow Na_2S_2O_3(aq)$

d. $2Fe(s) + 3Br_2(l) \rightarrow 2FeBr_3(s)$

e. $2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$

64.

a. $8Fe(s) + S_8(s) \rightarrow 8FeS(s)$

b. $4Co(s) + 3O_2(g) \rightarrow 2Co_2O_3(s)$

c. $Cl_2O_7(g) + H_2O(l) \rightarrow 2HClO_4(aq)$

65.

a. $CaSO_4(s) \rightarrow CaO(s) + SO_3(g)$

b. $Li_2CO_3(s) \rightarrow Li_2O(s) + CO_2(g)$

c. $2LiHCO_3(s) \rightarrow Li_2CO_3(s) + H_2O(g) + CO_2(g)$

d. $C_6H_6(l) \rightarrow 6C(s) + 3H_2(g)$

e. $4PBr_3(l) \rightarrow P_4(s) + 6Br_2(l)$

66.

a. $2NI_3(s) \rightarrow N_2(g) + 3I_2(s)$

b. $BaCO_3(s) \rightarrow BaO(s) + CO_2(g)$

c. $C_6H_{12}O_6(s) \rightarrow 6C(s) + 6H_2O(g)$

d. $Cu(NH_3)_4SO_4(s) \rightarrow CuSO_4(s) + 4NH_3(g)$

e. $3NaN_3(s) \rightarrow Na_3N(s) + 4N_2(g)$

67. A ***molecular equation*** uses the normal, uncharged formulas for the compounds involved. The ***complete ionic equation*** shows the compounds involved broken up into their respective ions (*all* ions present are shown). The ***net ionic equation*** shows only those ions that combine to form a precipitate, a gas, or a nonionic product such as water. The net ionic equation shows most clearly the species that are combining with each other.

d.

68. In several cases, the given ion may be precipitated by *many* reactants. The following are only three of the possible examples.

a. Chloride ion would precipitate when treated with solutions containing silver ion, lead(II) ion, or mercury(I) ion.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$
$$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$$
$$Hg_2^{2+}(aq) + 2Cl^-(aq) \rightarrow Hg_2Cl_2(s)$$

b. Calcium ion would precipitate when treated with solutions containing sulfate ion, carbonate ion, and phosphate ion.

$$Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$$
$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$
$$3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$$

c. Iron(III) ion would precipitate when treated with solutions containing hydroxide, sulfide, or carbonate ions.

$$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$$
$$2Fe^{3+}(aq) + 3S^{2-}(aq) \rightarrow Fe_2S_3(s)$$
$$2Fe^{3+}(aq) + 3CO_3^{2-}(aq) \rightarrow Fe_2(CO_3)_3(s)$$

d. Sulfate ion would precipitate when treated with solutions containing barium ion, calcium ion, or lead(II) ion.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$
$$Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$$
$$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$$

e. Mercury(I) ion would precipitate when treated with solutions containing chloride ion, sulfide ion, or carbonate ion.

$$Hg_2^{2+}(aq) + 2Cl^-(aq) \rightarrow Hg_2Cl_2(s)$$
$$Hg_2^{2+}(aq) + S^{2-}(aq) \rightarrow Hg_2S(s)$$
$$Hg_2^{2+}(aq) + CO_3^{2-}(aq) \rightarrow Hg_2CO_3(s)$$

f. Silver ion would precipitate when treated with solutions containing chloride ion, sulfide ion, or carbonate ion.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$
$$2Ag^+(aq) + S^{2-}(aq) \rightarrow Ag_2S(s)$$
$$2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$$

69.

a. $2Fe^{3+}(aq) + 3CO_3^{2-}(aq) \rightarrow Fe_2(CO_3)_3(s)$

b. $Hg_2^{2+}(aq) + 2\,Cl^-(aq) \rightarrow Hg_2Cl_2(s)$

c. no precipitate

d. $Cu^{2+}(aq) + S^{2-}(aq) \rightarrow CuS(s)$

e. $Pb^{2+}(aq) + 2Cl^{-}(aq) \rightarrow PbCl_2(s)$

f. $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

g. $Au^{3+}(aq) + 3OH^{-}(aq) \rightarrow Au(OH)_3(s)$

70. The formulas of the salts are indicated in boldface type.

a. $HNO_3(aq) + KOH(aq) \rightarrow H_2O(l) + \mathbf{KNO_3}(aq)$

b. $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + \mathbf{BaSO_4}(s)$

c. $HClO_4(aq) + NaOH(aq) \rightarrow H_2O(l) + \mathbf{NaClO_4}(aq)$

d. $2HCl(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + \mathbf{CaCl_2}(aq)$

71. For each cation, the precipitates that form with the anions listed in the right-hand column are given below. If no formula is listed, it should be assumed that the anion does *not* form a precipitate with the particular cation. See Table 7.1 for the Solubility Rules.

Ag^+ ion:	$AgCl$, Ag_2CO_3, $AgOH$, Ag_3PO_4, Ag_2S, Ag_2SO_4
Ba^{2+} ion:	$BaCO_3$, $Ba(OH)_2$, $Ba_3(PO_4)_2$, BaS, $BaSO_4$
Ca^{2+} ion:	$CaCO_3$, $Ca(OH)_2$, $Ca_3(PO_4)_2$, CaS, $CaSO_4$
Fe^{3+} ion:	$Fe_2(CO_3)_3$, $Fe(OH)_3$, $FePO_4$, Fe_2S_3
Hg_2^{2+} ion:	Hg_2Cl_2, Hg_2CO_3, $Hg_2(OH)_2$, $(Hg_2)_3(PO_4)_2$, Hg_2S
Na^+ ion:	All common salts are soluble.
Ni^{2+} ion:	$NiCO_3$, $Ni(OH)_2$, $Ni_3(PO_4)_2$, NiS
Pb^{2+} ion:	$PbCl_2$, $PbCO_3$, $Pb(OH)_2$, $Pb_3(PO_4)_2$, PbS, $PbSO_4$

72.

a. soluble (Rule 2: Most potassium salts are soluble.)

b. soluble (Rule 2: Most ammonium salts are soluble.)

c. insoluble (Rule 6: Most carbonate salts are only slightly soluble.)

d. insoluble (Rule 6: Most phosphate salts are only slightly soluble.)

e. soluble (Rule 2: Most sodium salts are soluble.)

f. insoluble (Rule 6: Most carbonate salts are only slightly soluble.)

g. soluble (Rule 3: Most chloride salts are soluble.)

73.

a. iron(III) hydroxide, $Fe(OH)_3$. Rule 5: Most hydroxide salts are only slightly soluble.

b. nickel(II) sulfide, NiS. Rule 6: Most sulfide salts are only slightly soluble.

c. silver chloride, $AgCl$. Rule 3: Although most chloride salts are soluble, $AgCl$ is a listed exception

d. barium carbonate, $BaCO_3$. Rule 6: Most carbonate salts are only slightly soluble.

e. mercury(I) chloride or mercurous chloride, Hg_2Cl_2. Rule 3: Although most chloride salts are soluble, Hg_2Cl_2 is a listed exception.

f. barium sulfate, $BaSO_4$. Rule 4: Although most sulfate salts are soluble, $BaSO_4$ is a listed exception.

74. The precipitates are marked in boldface type.

a. Rule 3: AgCl is listed as an exception.

$AgNO_3(aq) + HCl(aq) \rightarrow \mathbf{AgCl}(s) + HNO_3(aq)$

b. Rule 6: Most cabonate salts are only slightly soluble.

$CuSO_4(aq) + (NH_4)_2CO_3(aq) \rightarrow \mathbf{CuCO_3}(s) + (NH_4)_2SO_4(aq)$

c. Rule 6: Most carbonate salts are only slightly soluble.

$FeSO_4(aq) + K_2CO_3(aq) \rightarrow \mathbf{FeCO_3}(s) + K_2SO_4(aq)$

d. no reaction

e. Rule 6: Most carbonate salts are only slightly soluble.

$Pb(NO_3)_2(aq) + Li_2CO_3(aq) \rightarrow \mathbf{PbCO_3}(s) + 2LiNO_3(aq)$

f. Rule 5: Most hydroxide compounds are only slightly soluble.

$SnCl_4(aq) + 4NaOH(aq) \rightarrow \mathbf{Sn(OH)_4}(s) + 4NaCl(aq)$

75.

a. Rule 3: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

b. Rule 6: $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

c. Rule 3: $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

d. Rule 6: $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

76. $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$

$2Cr^{3+}(aq) + 3S^{2-}(aq) \rightarrow Cr_2S_3(s)$

$Ni^{2+}(aq) + S^{2-}(aq) \rightarrow NiS(s)$

77.

a. potassium hydroxide and perchloric acid

b. cesium hydroxide and nitric acid

c. potassium hydroxide and hydrochloric acid

d. sodium hydroxide and sulfuric acid

78. These anions tend to form insoluble precipitates with *many* metal ions. The following are illustrative for cobalt(II) chloride, tin(II) chloride, and copper(II) nitrate reacting with the sodium salts of the given anions.

a. $CoCl_2(aq) + Na_2S(aq) \rightarrow CoS(s) + 2NaCl(aq)$

$SnCl_2(aq) + Na_2S(aq) \rightarrow SnS(s) + 2NaCl(aq)$

$Cu(NO_3)_2(aq) + Na_2S(aq) \rightarrow CuS(s) + 2NaNO_3(aq)$

b. $CoCl_2(aq) + Na_2CO_3(aq) \rightarrow CoCO_3(s) + 2NaCl(aq)$

$SnCl_2(aq) + Na_2CO_3(aq) \rightarrow SnCO_3(s) + 2NaCl(aq)$

$Cu(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + 2NaNO_3(aq)$

c. $CoCl_2(aq) + 2NaOH(aq) \rightarrow Co(OH)_2(s) + 2NaCl(aq)$

$SnCl_2(aq) + 2NaOH(aq) \rightarrow Sn(OH)_2(s) + 2NaCl(aq)$

$Cu(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + 2NaNO_3(aq)$

d. $3CoCl_2(aq) + 2Na_3PO_4(aq) \rightarrow Co_3(PO_4)_2(s) + 6NaCl(aq)$

$3SnCl_2(aq) + 2Na_3PO_4(aq) \rightarrow Sn_3(PO_4)_2(s) + 6NaCl(aq)$

$3Cu(NO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Cu_3(PO_4)_2(s) + 6NaNO_3(aq)$

79. Fe_2S_3 is made up of Fe^{3+} and S^{2-} ions. Iron atoms each lose three electrons to become Fe^{3+} ions. Sulfur atoms each gain two electrons to become S^{2-} ions.

80.

a. $Na + O_2 \rightarrow Na_2O_2$

Balance sodium: $2Na + O_2 \rightarrow Na_2O_2$

Balanced equation: $2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$

b. $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

Equation is already balanced!

c. $Al_2O_3 \rightarrow Al + O_2$

Balance oxygen: $2Al_2O_3 \rightarrow Al + 3O_2$

Balance aluminum: $2Al_2O_3 \rightarrow 4Al + 3O_2$

Balanced equation: $2Al_2O_3(s) \rightarrow 4Al(s) + 3O_2(g)$

d. $Fe + Br_2 \rightarrow FeBr_3$

Balance bromine: $Fe + 3Br_2 \rightarrow 2FeBr_3$

Balance iron: $2Fe + 3Br_2 \rightarrow 2FeBr_3$

Balanced equation: $2Fe(s) + 3Br_2(l) \rightarrow 2FeBr_3(s)$

e. $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + H_2$

Balance nitrate ions: $Zn + 2HNO_3 \rightarrow Zn(NO_3)_2 + H_2$

Balanced equation: $Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$

81. For each reaction, the type of reaction is first identified, followed by some of the reasoning that leads to this choice (There may be more than one way in which you can recognize a particular type of reaction.).

a. oxidation–reduction (Mg changes from the elemental state to the combined state in $MgSO_4$; hydrogen changes from the combined to the elemental state.)

b. acid–base ($HClO_4$ is a strong acid, and RbOH is a strong base; water and a salt are produced.)

c. oxidation–reduction (Both Ca and O_2 change from the elemental to the combined state.)

d. acid–base (H_2SO_4 is a strong acid, and NaOH is a strong base; water and a salt are produced.)

e. precipitation (From the Solubility Rules of Table 7.1, $PbCO_3$ is insoluble.)

f. precipitation (From the Solubility Rules of Table 7.1, $CaSO_4$ is insoluble.)

g. acid–base (HNO_3 is a strong acid, and KOH is a strong base; water and a salt are produced.)

h. precipitation (From the Solubility Rules of Table 7.1, NiS is insoluble.)

i. oxidation–reduction (both Ni and Cl_2 change from the elemental to the combined state).

82.

a. $2C_4H_{10}(l) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

b. $C_4H_{10}O(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

c. $2C_4H_{10}O_2(l) + 11O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

83.

a. $4FeO(s) + O_2(g) \rightarrow 2Fe_2O_3(s)$

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

c. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

d. $16K(s) + S_8(s) \rightarrow 8K_2S(s)$

e. $6Na(s) + N_2(g) \rightarrow 2Na_3N(s)$

84.

a. $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

b. $2NaClO_3(s) \rightarrow 2NaCl(s) + 3O_2(g)$

c. $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

d. $C_{12}H_{22}O_{11}(s) \rightarrow 12C(s) + 11H_2O(g)$

e. $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$

85. For simplicity, the physical states of the substances are omitted.

$2Ba + O_2 \rightarrow 2BaO$

$Ba + S \rightarrow BaS$

$Ba + Cl_2 \rightarrow BaCl_2$

$3Ba + N_2 \rightarrow Ba_3N_2$

$Ba + Br_2 \rightarrow BaBr_2$

$4K + O_2 \rightarrow 2K_2O$

$2K + S \rightarrow K_2S$

$2K + Cl_2 \rightarrow 2KCl$

$6K + N_2 \rightarrow 2K_3N$

$2K + Br_2 \rightarrow 2KBr$

$2Mg + O_2 \rightarrow 2MgO$

$Mg + S \rightarrow MgS$

$Mg + Cl_2 \rightarrow MgCl_2$

$3Mg + N_2 \rightarrow Mg_3N_2$

$Mg + Br_2 \rightarrow MgBr_2$

$4Rb + O_2 \rightarrow 2Rb_2O$

$2Rb + S \rightarrow Rb_2S$

$2Rb + Cl_2 \rightarrow 2RbCl$

$6Rb + N_2 \rightarrow 2Rb_3N$

$2Rb + Br_2 \rightarrow 2RbBr$

$2Ca + O_2 \rightarrow 2CaO$

$Ca + S \rightarrow CaS$

$Ca + Cl_2 \rightarrow CaCl_2$

$3Ca + N_2 \rightarrow Ca_3N_2$

$Ca + Br_2 \rightarrow CaBr_2$

$4Li + O_2 \rightarrow 2Li_2$

$2Li + S \rightarrow Li_2S$

$2Li + Cl_2 \rightarrow 2LiCl$

$6Li + N_2 \rightarrow 2Li_3N$

$2Li + Br_2 \rightarrow 2LiBr$

86. $Al(s) + H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + H_2(g)$

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

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

$Co(s) + H_2SO_4(aq) \rightarrow CoSO_4(aq) + H_2(g)$

$Ni(s) + H_2SO_4(aq) \rightarrow NiSO_4(aq) + H_2(g)$

87. For simplicity, the physical states of the substances are omitted.

$Mg + Cl_2 \rightarrow MgCl_2$

$Ca + Cl_2 \rightarrow CaCl_2$

$Sr + Cl_2 \rightarrow SrCl_2$

$Ba + Cl_2 \rightarrow BaCl_2$

$Mg + Br_2 \rightarrow MgBr_2$

$Ca + Br_2 \rightarrow CaBr_2$

$Sr + Br_2 \rightarrow SrBr_2$

$Ba + Br_2 \rightarrow BaBr_2$

$2Mg + O_2 \rightarrow 2MgO$

$2Ca + O_2 \rightarrow 2CaO$

$2Sr + O_2 \rightarrow 2SrO$

$2Ba + O_2 \rightarrow 2BaO$

88.

a. one

b. one

c. two

d. two

e. three

89.

a. two; $O + 2e^- \rightarrow O^{2-}$

b. one; $F + e^- \rightarrow F^-$

c. three; $N + 3e^- \rightarrow N^{3-}$

d. one; $Cl + e^- \rightarrow Cl^-$

e. two; $S + 2e^- \rightarrow S^{2-}$

90. A very simple example that fits the bill is $C(s) + O_2(g) \rightarrow CO_2(g)$.

91.

a. $2I_4O_9(s) \rightarrow 2I_2O_6(s) + 2I_2(s) + 3O_2(g)$
oxidation–reduction, decomposition

b. $Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$
oxidation–reduction, single–displacement

c. $SiCl_4(l) + 2Mg(s) \rightarrow 2MgCl_2(s) + Si(s)$
oxidation–reduction, single–displacement

d. $CuCl_2(aq) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2AgCl(s)$
precipitation, double–displacement

e. $2Al(s) + 3Br_2(l) \rightarrow 2AlBr_3(s)$
oxidation–reduction, synthesis

92.

a. $2C_3H_8O(l) + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(g)$
oxidation–reduction, combustion

b. $HCl(aq) + AgC_2H_3O_2(aq) \rightarrow AgCl(s) + HC_2H_3O_2(aq)$
precipitation, double–displacement

c. $3HCl(aq) + Al(OH)_3(s) \rightarrow AlCl_3(aq) + 3H_2O(l)$
acid–base, double–displacement

d. $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
oxidation–reduction, decomposition

e. $N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$
oxidation–reduction, combustion

93. $2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$
$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
$2Fe(s) + O_2(g) \rightarrow 2FeO(s)$; $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$
$2Cr(s) + O_2(g) \rightarrow 2CrO(s)$; $4Cr(s) + 3O_2(g) \rightarrow 2Cr_2O_3(s)$
$2Ni(s) + O_2(g) \rightarrow 2NiO(s)$

94. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$
$Zn(s) + Cl_2(g) \rightarrow ZnCl_2(s)$
$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$
$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$; $Fe(s) + Cl_2(g) \rightarrow FeCl_2(s)$