▶ 실전 연습 해답

제2장

Self-Check Exercise 2.1

$357 = 3.57 \times 10^{2}$
$0.0055 = 5.5 \times 10^{-3}$

Self-Check Exercise 2.2

a. Three significant figures. The leading zeros (to the left of the 1) do not count, but the trailing zeros do.
b. Five significant figures. The one captive zero and the two trailing zeros all count.
c. This is an exact number obtained by counting the cars. It has an unlimited number of significant figures.

Self-Check Exercise 2.3

a. $12.6 \times 0.53 = 6.678 = 6.7$

b. $12.6 \times \underset{\text{Limiting}}{\overset{\text{Limiting}}{0.53}} = 6.7;$ $\begin{array}{r} 6.7 \;\text{Limiting} \\ \underline{-4.59} \\ 2.11 = 2.1 \end{array}$

c. $\begin{array}{r} 25.36 \\ \underline{-4.15} \\ 21.21 \end{array}$ $\dfrac{21.21}{2.317} = 9.15408 = 9.154$

Self-Check Exercise 2.4

$$0.750\ \cancel{\text{L}} \times \frac{1.06\ \text{qt}}{1\ \cancel{\text{L}}} = 0.795\ \text{qt}$$

Self-Check Exercise 2.5

$$225\ \frac{\cancel{\text{mi}}}{\text{h}} \times \frac{1760\ \cancel{\text{yd}}}{1\ \cancel{\text{mi}}} \times \frac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \frac{1\ \text{km}}{1000\ \cancel{\text{m}}} = 362\ \frac{\text{km}}{\text{h}}$$

Self-Check Exercise 2.6

The best way to solve this problem is to convert 172 K to Celsius degrees. To do this we will use the formula $T_{°C} = T_{K} - 273$.
In this case

$$T_{°C} = T_{K} - 273 = 172 - 273 = -101$$

So 172 K = −101 °C, which is a lower temperature than −75 °C. Thus 172 K is colder than −75 °C.

Self-Check Exercise 2.7

The problem is 41 °C = ? °F.
Using the formula

$$T_{°F} = 1.80(T_{°C}) + 32$$

we have

$$T_{°F} = ?\ °F = 1.80(41) + 32 = 74 + 32 = 106$$

That is, 41 °C = 106 °F.

Self-Check Exercise 2.8

This problem can be stated as 239 °F = ? °C.
Using the formula

$$T_{°C} = \frac{T_{°F} - 32}{1.80}$$

we have in this case

$$T_{°C} = ?\ °C = \frac{239 - 32}{1.80} = \frac{207}{1.80} = 115$$

That is, 239 °F = 115 °C.

Self-Check Exercise 2.9

We obtain the density of the cleaner by dividing its mass by its volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{28.1\ \text{g}}{35.8\ \text{mL}} = 0.785\ \text{g/mL}$$

This density identifies the liquid as isopropyl alcohol.

제3장

Self-Check Exercise 3.1

Items (a) and (c) are physical properties. When the solid gallium melts, it forms liquid gallium. There is no change in composition. Items (b) and (d) reflect the ability to change composition and are thus chemical properties. Statement (b) means that platinum does not react with oxygen to form some new substance. Statement (d) means that copper does react in the air to form a new substance, which is green.

Self-Check Exercise 3.2

a. Milk turns sour because new substances are formed. This is a chemical change.
b. Melting the wax is a physical change (a change of state). When the wax burns, new substances are formed. This is a chemical change.

Self-Check Exercise 3.3

a. Maple syrup is a homogeneous mixture of sugar and other substances dispersed uniformly in water.
b. Helium and oxygen form a homogeneous mixture.
c. Oil and vinegar salad dressing is a heterogeneous mixture. (Note the two distinct layers the next time you look at a bottle of dressing.)
d. Common salt is a pure substance (sodium chloride), so it always has the same composition. (Note that other substances such as iodine are often added to commercial preparations of table salt, which is mostly sodium chloride. Thus commercial table salt is a homogeneous mixture.)

제4장

Self-Check Exercise 4.1

a. P_4O_{10} b. UF_6 c. $AlCl_3$

Self-Check Exercise 4.2

In the symbol $^{90}_{38}\text{Sr}$, the number 38 is the atomic number, which represents the number of protons in the nucleus of a strontium atom. Because the atom is neutral overall, it must also have 38 electrons. The number 90 (the mass number) represents the number of protons plus the number of neutrons. Thus the number of neutrons is $A - Z = 90 - 38 = 52$.

Self-Check Exercise 4.3

The atom $^{201}_{80}\text{Hg}$ has 80 protons, 80 electrons, and 201 − 80 = 121 neutrons.

Self-Check Exercise 4.4

The atomic number for phosphorus is 15 and the mass number is 15 + 17 = 32. Thus the symbol for the atom is $^{32}_{15}\text{P}$.

Self-Check Exercise 4.5

Element	Symbol	Atomic Number	Metal or Nonmetal	Family Name
a. argon	Ar	18	nonmetal	noble gas
b. chlorine	Cl	17	nonmetal	halogen
c. barium	Ba	56	metal	alkaline earth metal
d. cesium	Cs	55	metal	alkali metal

Self-Check Exercise 4.6

a. KI $(1+) + (1-) = 0$
b. Mg_3N_2 $3(2+) + 2(3-) = (6+) + (6-) = 0$
c. Al_2O_3 $2(3+) + 3(2-) = 0$

제 5 장

Self-Check Exercise 5.1

a. rubidium oxide
b. strontium iodide
c. potassium sulfide

Self-Check Exercise 5.2

a. The compound $PbBr_2$ must contain Pb^{2+}—named lead(II)—to balance the charges of the two Br^- ions. Thus the name is lead(II) bromide. The compound $PbBr_4$ must contain Pb^{4+}—named lead(IV)—to balance the charges of the four Br^- ions. The name is therefore lead(IV) bromide.
b. The compound FeS contains the S^{2-} ion (sulfide) and thus the iron cation present must be Fe^{2+}, iron(II). The name is iron(II) sulfide. The compound Fe_2S_3 contains three S^{2-} ions and two iron cations of unknown charge. We can determine the iron charge from the following:

$$2(?+) + 3(2-) = 0$$

Iron charge S^{2-} charge

In this case, ? must represent 3 because

$$2(3+) + 3(2-) = 0$$

Thus Fe_2S_3 contains Fe^{3+} and S^{2-}, and its name is iron(III) sulfide.
c. The compound $AlBr_3$ contains Al^{3+} and Br^-. Because aluminum forms only one ion (Al^{3+}), no Roman numeral is required. The name is aluminum bromide.
d. The compound Na_2S contains Na^+ and S^{2-} ions. The name is sodium sulfide. (Because sodium forms only Na^+, no Roman numeral is needed.)
e. The compound $CoCl_3$ contains three Cl^- ions. Thus the cobalt cation must be Co^{3+}, which is named cobalt(III) because cobalt is a transition metal and can form more than one type of cation. Thus the name of $CoCl_3$ is cobalt(III) chloride.

Self-Check Exercise 5.3

Compound	Individual Names	Prefixes	Name
a. CCl_4	carbon chloride	none *tetra-*	carbon tetrachloride
b. NO_2	nitrogen oxide	none *di-*	nitrogen dioxide
c. IF_5	iodine fluoride	none *penta-*	iodine pentafluoride

Self-Check Exercise 5.4

a. silicon dioxide
b. dioxygen difluoride
c. xenon hexafluoride

Self-Check Exercise 5.5

a. chlorine trifluoride
b. vanadium(V) fluoride
c. copper(I) chloride
d. manganese(IV) oxide
e. magnesium oxide
f. water

Self-Check Exercise 5.6

a. calcium hydroxide
b. sodium phosphate
c. potassium permanganate
d. ammonium dichromate
e. cobalt(II) perchlorate (Perchlorate has a 1− charge, so the cation must be Co^{2+} to balance the two ClO_4^- ions.)
f. potassium chlorate
g. copper(II) nitrite (This compound contains two NO_2^- (nitrite) ions and thus must contain a Cu^{2+} cation.)

Self-Check Exercise 5.7

Compound **Name**

a. $NaHCO_3$ sodium hydrogen carbonate
Contains Na^+ and HCO_3^-; often called sodium bicarbonate (common name).
b. $BaSO_4$ barium sulfate
Contains Ba^{2+} and SO_4^{2-}.
c. $CsClO_4$ cesium perchlorate
Contains Cs^+ and ClO_4^-.
d. BrF_5 bromine pentafluoride
Both nonmetals (Type III binary).
e. NaBr sodium bromide
Contains Na^+ and Br^- (Type I binary).
f. KOCl potassium hypochlorite
Contains K^+ and OCl^-.
g. $Zn_3(PO_4)_2$ zinc(II) phosphate
Contains Zn^{2+} and PO_4^{3-}; Zn is a transition metal and officially requires a Roman numeral. However, because Zn forms only the Zn^{2+} cation, the II is usually left out. Thus the name of the compound is usually given as zinc phosphate.

Self-Check Exercise 5.8

Name **Chemical Formula**

a. ammonium sulfate $(NH_4)_2SO_4$
Two ammonium ions (NH_4^+) are required for each sulfate ion (SO_4^{2-}) to achieve charge balance.
b. vanadium(V) fluoride VF_5
The compound contains V^{5+} ions and requires five F^- ions for charge balance.
c. disulfur dichloride S_2Cl_2
The prefix *di-* indicates two of each atom.
d. rubidium peroxide Rb_2O_2
Because rubidium is in Group 1, it forms only 1+ ions. Thus two Rb^+ ions are needed to balance the 2− charge on the peroxide ion (O_2^{2-}).
e. aluminum oxide Al_2O_3
Aluminum forms only 3+ ions. Two Al^{3+} ions are required to balance the charge on three O^{2-} ions.

제 6 장

Self-Check Exercise 6.1

a. $Mg(s) + H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)$
Note that magnesium (which is in Group 2) always forms the Mg^{2+} cation and thus requires two OH^- anions for a zero net charge.

b. Ammonium dichromate contains the polyatomic ions NH_4^+ and $Cr_2O_7^{2-}$ (you should have these memorized). Because NH_4^+ has a 1+ charge, two NH_4^+ cations are required for each $Cr_2O_7^{2-}$, with it 2− charge, to give the formula $(NH_4)_2Cr_2O_7$. Chromium(III) oxide contains Cr^{3+} ions—signified by chromium(III)—and O^{2-} (the oxide ion). To achieve a net charge of zero, the solid must contain two Cr^{3+} ions for every three O^{2-} ions, so the formula is Cr_2O_3. Nitrogen gas contains diatomic molecules and is written $N_2(g)$, and gaseous water is written $H_2O(g)$. Thus the unbalanced equation for the decomposition of ammonium dichromate is

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

c. Gaseous ammonia, $NH_3(g)$, and gaseous oxygen, $O_2(g)$, react to form nitrogen monoxide gas, $NO(g)$, plus gaseous water, $H_2O(g)$. The unbalanced equation is

$$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$$

Self-Check Exercise 6.2

Step 1 The reactants are propane, $C_3H_8(g)$, and oxygen, $O_2(g)$; the products are carbon dioxide, $CO_2(g)$, and water, $H_2O(g)$. All are in the gaseous state.

Step 2 The unbalanced equation for the reaction is

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

Step 3 We start with C_3H_8 because it is the most complicated molecule. C_3H_8 contains three carbon atoms per molecule, so a coefficient of 3 is needed for CO_2.

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

Also, each C_3H_8 molecule contains eight hydrogen atoms, so a coefficient of 4 is required for H_2O.

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

The final element to be balanced is oxygen. Note that the left side of the equation now has two oxygen atoms, and the right side has ten. We can balance the oxygen by using a coefficient of 5 for O_2.

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

Step 4 ***Check:***

$$\underset{\text{Reactant atoms}}{3\text{ C, 8 H, 10 O}} \rightarrow \underset{\text{Product atoms}}{3\text{ C, 8 H, 10 O}}$$

We cannot divide all coefficients by a given integer to give smaller integer coefficients.

Self-Check Exercise 6.3

a. $NH_4NO_2(s) \rightarrow N_2(g) + H_2O(g)$ (unbalanced)
$NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(g)$ (balanced)

b. $NO(g) \rightarrow N_2O(g) + NO_2(g)$ (unbalanced)
$3NO(g) \rightarrow N_2O(g) + NO_2(g)$ (balanced)

c. $HNO_3(l) \rightarrow NO_2(g) + H_2O(l) + O_2(g)$ (unbalanced)
$4HNO_3(l) \rightarrow 4NO_2(g) + 2H_2O(l) + O_2(g)$ (balanced)

제 7 장

Self-Check Exercise 7.1

a. The ions present are

$$\underbrace{Ba^{2+}(aq) + 2NO_3^-(aq)}_{\text{Ions in } Ba(NO_3)_2(aq)} + \underbrace{Na^+(aq) + Cl^-(aq)}_{\text{Ions in } NaCl(aq)} \rightarrow$$

Exchanging the anions gives the possible solid products $BaCl_2$ and $NaNO_3$. Using Table 7.1, we see that both substances are very soluble (rules 1, 2, and 3). Thus no solid forms.

b. The ions present in the mixed solution before any reaction occurs are

$$\underbrace{2Na^+(aq) + S^{2-}(aq)}_{\text{Ions in } Na_2S(aq)} + \underbrace{Cu^{2+}(aq) + 2NO_3^-(aq)}_{\text{Ions in } Cu(NO_3)_2(aq)} \rightarrow$$

Exchanging the anions gives the possible solid products CuS and $NaNO_3$. According to rules 1 and 2 in Table 7.1, $NaNO_3$ is soluble, and by rule 6, CuS should be insoluble. Thus CuS will precipitate. The balanced equation is

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

c. The ions present are

$$\underbrace{NH_4^+(aq) + Cl^-(aq)}_{\text{Ions in } NH_4Cl(aq)} + \underbrace{Pb^{2+}(aq) + 2NO_3^-(aq)}_{\text{Ions in } Pb(NO_3)_2(aq)} \rightarrow$$

Exchanging the anions gives the possible solid products NH_4NO_3 and $PbCl_2$. NH_4NO_3 is soluble (rules 1 and 2) and $PbCl_2$ is insoluble (rule 3). Thus $PbCl_2$ will precipitate. The balanced equation is

$$2NH_4Cl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2NH_4NO_3(aq)$$

Self-Check Exercise 7.2

a. *Molecular equation:*
$Na_2S(aq) + Cu(NO_3)_2(aq) \rightarrow CuS(s) + 2NaNO_3(aq)$
Complete ionic equation:
$2Na^+(aq) + S^{2-}(aq) + Cu^{2+}(aq) + 2NO_3^-(aq) \rightarrow CuS(s) + 2Na^+(aq) + 2NO_3^-(aq)$
Net ionic equation:
$S^{2-}(aq) + Cu^{2+}(aq) \rightarrow CuS(s)$

b. *Molecular equation:*
$2NH_4Cl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2NH_4NO_3(aq)$
Complete ionic equation:
$2NH_4^+(aq) + 2Cl^-(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \rightarrow PbCl_2(s) + 2NH_4^+(aq) + 2NO_3^-(aq)$
Net ionic equation:
$2Cl^-(aq) + Pb^{2+}(aq) \rightarrow PbCl_2(s)$

Self-Check Exercise 7.3

a. The compound NaBr contains the ions Na^+ and Br^-. Thus each sodium atom loses one electron ($Na \rightarrow Na^+ + e^-$), and each bromine atom gains one electron ($Br + e^- \rightarrow Br^-$).

$$Na + Na + Br - Br \rightarrow (Na^+Br^-) + (Na^+Br^-)$$
$$e^- \quad e^-$$

b. The compound CaO contains the Ca^{2+} and O^{2-} ions. Thus each calcium atom loses two electrons ($Ca \rightarrow Ca^{2+} + 2e^-$), and each oxygen atom gains two electrons ($O + 2e^- \rightarrow O^{2-}$).

$$Ca + Ca + O - O \rightarrow (Ca^{2+}O^{2-}) + (Ca^{2+}O^{2-})$$
$$2e^- \quad 2e^-$$

Self-Check Exercise 7.4

a. oxidation–reduction reaction; combustion reaction
b. synthesis reaction; oxidation–reduction reaction; combustion reaction
c. synthesis reaction; oxidation–reduction reaction
d. decomposition reaction; oxidation–reduction reaction
e. precipitation reaction (and double displacement)
f. synthesis reaction; oxidation–reduction reaction
g. acid–base reaction (and double displacement)
h. combustion reaction; oxidation–reduction reaction

제8장

Self-Check Exercise 8.1

The average mass of nitrogen is 14.01 amu. The appropriate equivalence statement is 1 N atom = 14.01 amu, which yields the conversion factor we need:

$$23 \ \cancel{\text{N atoms}} \times \frac{14.01 \text{ amu}}{\cancel{\text{N atom}}} = 322.2 \text{ amu}$$

(exact)

Self-Check Exercise 8.2

The average mass of oxygen is 16.00 amu, which gives the equivalence statement 1 O atom = 16.00 amu. The number of oxygen atoms present is

$$288 \ \cancel{\text{amu}} \times \frac{1 \text{ O atom}}{16.00 \ \cancel{\text{amu}}} = 18.0 \text{ O atoms}$$

Self-Check Exercise 8.3

Note that the sample of 5.00×10^{20} atoms of chromium is less than 1 mole (6.022×10^{23} atoms) of chromium. What fraction of a mole it represents can be determined as follows:

$$5.00 \times 10^{20} \ \cancel{\text{atoms Cr}} \times \frac{1 \text{ mol Cr}}{6.022 \times 10^{23} \ \cancel{\text{atoms Cr}}} = 8.30 \times 10^{-4} \text{ mol Cr}$$

Because the mass of 1 mole of chromium atoms is 52.00 g, the mass of 5.00×10^{20} atoms can be determined as follows:

$$8.30 \times 10^{-4} \ \cancel{\text{mol Cr}} \times \frac{52.00 \text{ g Cr}}{1 \ \cancel{\text{mol Cr}}} = 4.32 \times 10^{-2} \text{ g Cr}$$

Self-Check Exercise 8.4

Each molecule of C_2H_3Cl contains two carbon atoms, three hydrogen atoms, and one chlorine atom, so 1 mole of C_2H_3Cl molecules contains 2 moles of C atoms, 3 moles of H atoms, and 1 mole of Cl atoms.

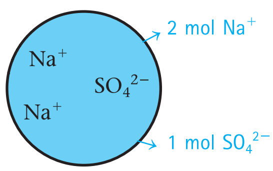
Mass of 2 mol C atoms:	$2 \times 12.01 =$	24.02 g
Mass of 3 mol H atoms:	$3 \times 1.008 =$	3.024 g
Mass of 1 mol Cl atoms:	$1 \times 35.45 =$	35.45 g
		62.494 g

The molar mass of C_2H_3Cl is 62.49 g (rounding to the correct number of significant figures).

Self-Check Exercise 8.5

The formula for sodium sulfate is Na_2SO_4. One mole of Na_2SO_4 contains 2 moles of sodium ions and 1 mole of sulfate ions.

1 mole of $Na_2SO_4 \rightarrow$ 1 mole of



Mass of 2 mol Na^+	$= 2 \times 22.99$	= 45.98 g
Mass of 1 mol SO_4^{2-}	$= 32.07 + 4(16.00)$	= 96.07 g
Mass of 1 mol Na_2SO_4		= 142.05 g

The molar mass for sodium sulfate is 142.05 g.

A sample of sodium sulfate with a mass of 300.0 g represents more than 1 mol. (Compare 300.0 g to the molar mass of Na_2SO_4.) We calculate the number of moles of Na_2SO_4 present in 300.0 g as follows:

$$300.0 \ \cancel{\text{g } Na_2SO_4} \times \frac{1 \text{ mol } Na_2SO_4}{142.05 \ \cancel{\text{g } Na_2SO_4}} = 2.112 \text{ mol } Na_2SO_4$$

Self-Check Exercise 8.6

First we must compute the mass of 1 mole of C_2F_4 molecules (the molar mass). Because 1 mole of C_2F_4 contains 2 moles of C atoms and 4 moles of F atoms, we have

$$2 \ \cancel{\text{mol}} \text{ C} \times \frac{12.01 \text{ g}}{\cancel{\text{mol}}} = 24.02 \text{ g C}$$

$$4 \ \cancel{\text{mol}} \text{ F} \times \frac{19.00 \text{ g}}{\cancel{\text{mol}}} = 76.00 \text{ g F}$$

Mass of 1 mole of C_2F_4: 100.02 g = molar mass

Using the equivalence statement 100.02 g C_2F_4 = 1 mole C_2F_4, we calculate the moles of C_2F_4 units in 135 g of Teflon.

$$135 \ \cancel{\text{g } C_2F_4} \text{ units} \times \frac{1 \text{ mol } C_2F_4}{100.02 \ \cancel{\text{g } C_2F_4}} = 1.35 \text{ mol } C_2F_4 \text{ units}$$

Next, using the equivalence statement 1 mol = 6.022×10^{23} units, we calculate the number of C_2F_4 units in 135 mol of Teflon.

$$135 \ \cancel{\text{mol}} \ C_2F_4 \times \frac{6.022 \times 10^{23} \text{ units}}{1 \ \cancel{\text{mol}}} = 8.13 \times 10^{23} \ C_2F_4 \text{ units}$$

Self-Check Exercise 8.7

The molar mass of penicillin F is computed as follows:

$$\text{C: } 14 \ \cancel{\text{mol}} \times 12.01 \frac{\text{g}}{\cancel{\text{mol}}} = 168.1 \text{ g}$$

$$\text{H: } 20 \ \cancel{\text{mol}} \times 1.008 \frac{\text{g}}{\cancel{\text{mol}}} = 20.16 \text{ g}$$

$$\text{N: } 2 \ \cancel{\text{mol}} \times 14.01 \frac{\text{g}}{\cancel{\text{mol}}} = 28.02 \text{ g}$$

$$\text{S: } 1 \ \cancel{\text{mol}} \times 32.07 \frac{\text{g}}{\cancel{\text{mol}}} = 32.07 \text{ g}$$

$$\text{O: } 4 \ \cancel{\text{mol}} \times 16.00 \frac{\text{g}}{\cancel{\text{mol}}} = 64.00 \text{ g}$$

Mass of 1 mole of $C_{14}H_{20}N_2SO_4$ = 312.39 g = 312.4 g

$$\text{Mass percent of C} = \frac{168.1 \text{ g C}}{312.4 \text{ g } C_{14}H_{20}N_2SO_4} \times 100\% = 53.81\%$$

$$\text{Mass percent of H} = \frac{20.16 \text{ g H}}{312.4 \text{ g } C_{14}H_{20}N_2SO_4} \times 100\% = 6.453\%$$

$$\text{Mass percent of N} = \frac{28.02 \text{ g N}}{312.4 \text{ g } C_{14}H_{20}N_2SO_4} \times 100\% = 8.969\%$$

$$\text{Mass percent of S} = \frac{32.07 \text{ g S}}{312.4 \text{ g } C_{14}H_{20}N_2SO_4} \times 100\% = 10.27\%$$

$$\text{Mass percent of O} = \frac{64.00 \text{ g O}}{312.4 \text{ g } C_{14}H_{20}N_2SO_4} \times 100\% = 20.49\%$$

Check: The percentages add up to 99.99%.

Self-Check Exercise 8.8

Step 1 0.6884 g lead and 0.2356 g chlorine

Step 2

$$0.6884 \ \cancel{\text{g Pb}} \times \frac{1 \text{ mol Pb}}{207.2 \ \cancel{\text{g Pb}}} = 0.003322 \text{ mol Pb}$$

$$0.2356 \ \cancel{\text{g Cl}} \times \frac{1 \text{ mol Cl}}{35.45 \ \cancel{\text{g Cl}}} = 0.006646 \text{ mol Cl}$$

Step 3 $\dfrac{0.003322 \text{ mol Pb}}{0.003322} = 1.000 \text{ mol Pb}$

$\dfrac{0.006646 \text{ mol Cl}}{0.003322} = 2.001 \text{ mol Cl}$

These numbers are very close to integers, so step 4 is unnecessary. The empirical formula is $PbCl_2$.

Self-Check Exercise 8.9

Step 1 0.8007 g C, 0.9333 g N, 0.2016 g H, and 2.133 g O

Step 2 $0.8007 \cancel{\text{g C}} \times \dfrac{1 \text{ mol C}}{12.01 \cancel{\text{g C}}} = 0.06667 \text{ mol C}$

$0.9333 \cancel{\text{g N}} \times \dfrac{1 \text{ mol N}}{14.01 \cancel{\text{g N}}} = 0.06662 \text{ mol N}$

$0.2016 \cancel{\text{g H}} \times \dfrac{1 \text{ mol H}}{1.008 \cancel{\text{g H}}} = 0.2000 \text{ mol H}$

$2.133 \cancel{\text{g O}} \times \dfrac{1 \text{ mol O}}{16.00 \cancel{\text{g O}}} = 0.1333 \text{ mol O}$

Step 3 $\dfrac{0.06667 \text{ mol C}}{0.06667} = 1.001 \text{ mol C}$

$\dfrac{0.06662 \text{ mol N}}{0.06667} = 1.000 \text{ mol N}$

$\dfrac{0.2000 \text{ mol H}}{0.06662} = 3.002 \text{ mol H}$

$\dfrac{0.1333 \text{ mol O}}{0.06662} = 2.001 \text{ mol O}$

The empirical formula is CNH_3O_2.

Self-Check Exercise 8.10

Step 1 In 100.00 g of Nylon-6 the masses of elements present are 63.68 g C, 12.38 g N, 9.80 g H, and 14.14 g O.

Step 2 $63.68 \cancel{\text{g C}} \times \dfrac{1 \text{ mol C}}{12.01 \cancel{\text{g C}}} = 5.302 \text{ mol C}$

$12.38 \cancel{\text{g N}} \times \dfrac{1 \text{ mol N}}{14.01 \cancel{\text{g N}}} = 0.8837 \text{ mol N}$

$9.80 \cancel{\text{g H}} \times \dfrac{1 \text{ mol H}}{1.008 \cancel{\text{g H}}} = 9.72 \text{ mol H}$

$14.14 \cancel{\text{g O}} \times \dfrac{1 \text{ mol O}}{16.00 \cancel{\text{g O}}} = 0.8838 \text{ mol O}$

Step 3 $\dfrac{5.302 \text{ mol C}}{0.8836} = 6.000 \text{ mol C}$

$\dfrac{0.8837 \text{ mol N}}{0.8837} = 1.000 \text{ mol N}$

$\dfrac{9.72 \text{ mol H}}{0.8837} = 11.0 \text{ mol H}$

$\dfrac{0.8838 \text{ mol O}}{0.8837} = 1.000 \text{ mol O}$

The empirical formula for Nylon-6 is $C_6NH_{11}O$.

Self-Check Exercise 8.11

Step 1 First we convert the mass percents to mass in grams. In 100.0 g of the compound, there are 71.65 g of chlorine, 24.27 g of carbon, and 4.07 g of hydrogen.

Step 2 We use these masses to compute the moles of atoms present.

$$71.65 \cancel{\text{g Cl}} \times \frac{1 \text{ mol Cl}}{35.45 \cancel{\text{g Cl}}} = 2.021 \text{ mol Cl}$$

$$24.27 \cancel{\text{g C}} \times \frac{1 \text{ mol C}}{12.01 \cancel{\text{g C}}} = 2.021 \text{ mol C}$$

$$4.07 \cancel{\text{g H}} \times \frac{1 \text{ mol H}}{1.008 \cancel{\text{g H}}} = 4.04 \text{ mol H}$$

Step 3 Dividing each mole value by 2.021 (the smallest number of moles present), we obtain the empirical formula $ClCH_2$.

To determine the molecular formula, we must compare the empirical formula mass to the molar mass. The empirical formula mass is 49.48.

Cl: 35.45
C: 12.01
2 H: $2 \times (1.008)$
$ClCH_2$: 49.48 = empirical formula mass

The molar mass is known to be 98.96. We know that

$$\text{Molar mass} = n \times (\text{empirical formula mass})$$

So we can obtain the value of n as follows:

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96}{49.48} = 2$$

$$\text{Molecular formula} = (ClCH_2)_2 = Cl_2C_2H_4$$

This substance is composed of molecules with the formula $Cl_2C_2H_4$.

제9장

Self-Check Exercise 9.1

The problem can be stated as follows:

$$4.30 \text{ mol } C_3H_8 \xrightarrow[\text{yields}]{} ? \text{ mol } CO_2$$

From the balanced equation

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

we derive the equivalence statement

$$1 \text{ mol } C_3H_8 = 3 \text{ mol } CO_2$$

The appropriate conversion factor (moles of C_3H_8 must cancel) is $3 \text{ mol } CO_2 / 1 \text{ mol } C_3H_8$, and the calculation is

$$4.30 \cancel{\text{mol } C_3H_8} \times \frac{3 \text{ mol } CO_2}{1 \cancel{\text{mol } C_3H_8}} = 12.9 \text{ mol } CO_2$$

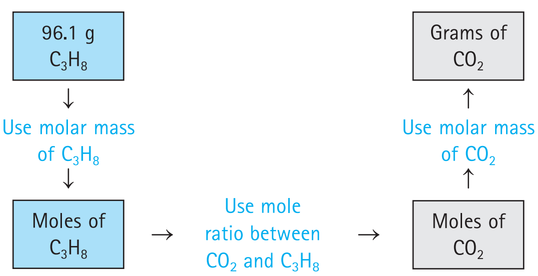
Thus we can say

$$4.30 \text{ mol } C_3H_8 \text{ yields } 12.9 \text{ mol } CO_2$$

Self-Check Exercise 9.2

The problem can be sketched as follows:

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



We have already done the first step in Example 9.4.

To find out how many moles of CO_2 can be produced from 2.18 moles of C_3H_8, we see from the balanced equation that 3 moles of CO_2 is produced for each mole of C_3H_8 reacted. The mole ratio we need is 3 mol CO_2/1 mol C_3H_8. The conversion is therefore

$$2.18\ \cancel{\text{mol } C_3H_8} \times \frac{3\text{ mol } CO_2}{1\ \cancel{\text{mol } C_3H_8}} = 6.54\text{ mol } CO_2$$

Next, using the molar mass of CO_2, which is 12.01 + 32.00 = 44.01 g, we calculate the mass of CO_2 produced.

$$6.54\ \cancel{\text{mol } CO_2} \times \frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 288\text{ g } CO_2$$

The sequence of steps we took to find the mass of carbon dioxide produced from 96.1 g of propane is summarized in the following diagram.

96.1 g C_3H_8	→	$\frac{1\text{ mol } C_3H_8}{44.09\text{ g } C_3H_8}$	→	2.18 mol C_3H_8
2.18 mol C_3H_8	→	$\frac{3\text{ mol } CO_2}{1\text{ mol } C_3H_8}$	→	6.54 mol CO_2
6.54 mol CO_2	→	$\frac{44.01\text{ g } CO_2}{1\text{ mol } CO_2}$	→	288 g CO_2
Mass				Moles

Self-Check Exercise 9.3

We sketch the problem as follows:

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

96.1 g C_3H_8				Grams of H_2O
↓ $\frac{1\text{ mol } C_3H_8}{44.09\text{ g}}$ ↓				↑ $\frac{18.02\text{ g}}{\text{mol } H_2O}$ ↑
Moles of C_3H_8	→	$\frac{4\text{ mol } H_2O}{1\text{ mol } C_3H_8}$	→	Moles of H_2O

Then we do the calculations.

96.1 g C_3H_8	→	$\frac{1\text{ mol } C_3H_8}{44.09\text{ g}}$	→	2.18 mol C_3H_8
2.18 mol C_3H_8	→	$\frac{4\text{ mol } H_2O}{1\text{ mol } C_3H_8}$	→	8.72 mol H_2O
8.72 mol H_2O	→	$\frac{18.02\text{ g}}{\text{mol } H_2O}$	→	157 g H_2O

Therefore, 157 g of H_2O is produced from 96.1 g C_3H_8.

Self-Check Exercise 9.4

a. We first write the balanced equation.

$$SiO_2(s) + 4HF(aq) \rightarrow SiF_4(g) + 2H_2O(l)$$

The map of the steps required is

$$SiO_2(s) + 4HF(aq) \rightarrow SiF_4(g) + 2H_2O(l)$$

5.68 g SiO_2				Grams of HF
↓ Use molar mass of SiO_2 ↓				↑ Use molar mass of HF ↑
Moles of SiO_2	→	Use mole ratio between HF and SiO_2	→	Moles of HF

We convert 5.68 g of SiO_2 to moles as follows:

$$5.68\text{ g } \cancel{SiO_2} \times \frac{1\text{ mol } SiO_2}{60.09\text{ g } \cancel{SiO_2}} = 9.45 \times 10^{-2}\text{ mol } SiO_2$$

Using the balanced equation, we obtain the appropriate mole ratio and convert to moles of HF.

$$9.45 \times 10^{-2}\ \cancel{\text{mol } SiO_2} \times \frac{4\text{ mol HF}}{1\ \cancel{\text{mol } SiO_2}} = 3.78 \times 10^{-1}\text{ mol HF}$$

Finally, we calculate the mass of HF by using its molar mass.

$$3.78 \times 10^{-1}\ \cancel{\text{mol HF}} \times \frac{20.01\text{ g HF}}{\cancel{\text{mol HF}}} = 7.56\text{ g HF}$$

b. The map for this problem is

$$SiO_2(s) + 4HF(aq) \rightarrow SiF_4(g) + 2H_2O(l)$$

5.68 g SiO_2				Grams of H_2O
↓ Use molar mass of SiO_2 ↓				↑ Use molar mass of H_2O ↑
Moles of SiO_2	→	Use mole ratio between H_2O and SiO_2	→	Moles of H_2O

We have already accomplished the first conversion in part a. Using the balanced equation, we obtain moles of H_2O as follows:

$$9.45 \times 10^{-2}\ \cancel{\text{mol } SiO_2} \times \frac{2\text{ mol } H_2O}{1\ \cancel{\text{mol } SiO_2}} = 1.89 \times 10^{-1}\text{ mol } H_2O$$

The mass of water formed is

$$1.89 \times 10^{-1}\ \cancel{\text{mol } H_2O} \times \frac{18.02\text{ g } H_2O}{\cancel{\text{mol } H_2O}} = 3.41\text{ g } H_2O$$

Self-Check Exercise 9.5

In this problem, we know the mass of the product to be formed by the reaction

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$

and we want to find the masses of reactants needed. The procedure is the same one we have been following. We must first convert the mass of CH_3OH to moles, then use the balanced equation to obtain moles of H_2 and CO needed, and then convert these moles to masses. Using the molar mass of CH_3OH (32.04 g/mol), we convert to moles of CH_3OH.

First we convert kilograms to grams.

$$6.0\ \cancel{\text{kg}}\ CH_3OH \times \frac{1000\text{ g}}{\cancel{\text{kg}}} = 6.0 \times 10^3\text{ g } CH_3OH$$

Next we convert 6.0×10^3 g CH_3OH to moles of CH_3OH, using the conversion factor 1 mol CH_3OH/32.04 g CH_3OH.

$$6.0 \times 10^3\text{ g } \cancel{CH_3OH} \times \frac{1\text{ mol } CH_3OH}{32.04\text{ g } \cancel{CH_3OH}} = 1.9 \times 10^2\text{ mol } CH_3OH$$

Then we have two questions to answer:

$$\boxed{\text{? mol of } H_2} \xrightarrow[\text{to produce}]{\text{required}} \boxed{1.9 \times 10^2 \text{ mol } CH_3OH}$$

$$\boxed{\text{? mol of } CO} \xrightarrow[\text{to produce}]{\text{required}} \boxed{1.9 \times 10^2 \text{ mol } CH_3OH}$$

To answer these questions, we use the balanced equation

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$

to obtain mole ratios between the reactants and the products. In the balanced equation the coefficients for both CO and CH_3OH are 1, so we can write the equivalence statement

$$1 \text{ mol CO} = 1 \text{ mol } CH_3OH$$

Using the mole ratio 1 mol CO/1 mol CH_3OH, we can now convert from moles of CH_3OH to moles of CO.

$$1.9 \times 10^2 \cancel{\text{mol } CH_3OH} \times \frac{1 \text{ mol CO}}{1 \cancel{\text{mol } CH_3OH}} = 1.9 \times 10^2 \text{ mol CO}$$

To calculate the moles of H_2 required, we construct the equivalence statement between CH_3OH and H_2, using the coefficients in the balanced equation.

$$2 \text{ mol } H_2 = 1 \text{ mol } CH_3OH$$

Using the mole ratio 2 mol H_2/1 mol CH_3OH, we can convert moles of CH_3OH to moles of H_2.

$$1.9 \times 10^2 \cancel{\text{mol } CH_3OH} \times \frac{2 \text{ mol } H_2}{1 \cancel{\text{mol } CH_3OH}} = 3.8 \times 10^2 \text{ mol } H_2$$

We now have the moles of reactants required to produce 6.0 kg of CH_3OH. Since we need the masses of reactants, we must use the molar masses to convert from moles to mass.

$$1.9 \times 10^2 \cancel{\text{mol CO}} \times \frac{28.01 \text{ g CO}}{1 \cancel{\text{mol CO}}} = 5.3 \times 10^3 \text{ g CO}$$

$$3.8 \times 10^2 \cancel{\text{mol } H_2} \times \frac{2.016 \text{ g } H_2}{1 \cancel{\text{mol } H_2}} = 7.7 \times 10^2 \ H_2$$

Therefore, we need 5.3×10^3 g CO to react with 7.7×10^2 g H_2 to form 6.0×10^3 g (6.0 kg) of CH_3OH. This whole process is mapped in the following diagram.

$$\begin{tikzcd}
 & \boxed{6.0 \times 10^3 \text{ g } CH_3OH} \arrow[d, "\frac{1 \text{ mol } CH_3OH}{32.04 \text{ g } CH_3OH}"] & \\
 & \boxed{1.9 \times 10^2 \text{ mol } CH_3OH} \arrow[dl, "\frac{1 \text{ mol CO}}{1 \text{ mol } CH_3OH}"'] \arrow[dr, "\frac{2 \text{ mol } H_2}{1 \text{ mol } CH_3OH}"] & \\
\boxed{3.8 \times 10^2 \text{ mol } H_2} \arrow[d, "\frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}"] & & \boxed{1.9 \times 10^2 \text{ mol CO}} \arrow[d, "\frac{28.01 \text{ g CO}}{1 \text{ mol CO}}"] \\
\boxed{7.7 \times 10^2 \text{ g } H_2} & & \boxed{5.3 \times 10^3 \text{ g CO}}
\end{tikzcd}$$

Self-Check Exercise 9.6

Step 1 The balanced equation for the reaction is

$$6Li(s) + N_2(g) \rightarrow 2Li_3N(s)$$

Step 2 To determine the limiting reactant, we must convert the masses of lithium (atomic mass = 6.941 g) and nitrogen (molar mass = 28.02 g) to moles.

$$56.0 \cancel{\text{g Li}} \times \frac{1 \text{ mol Li}}{6.941 \cancel{\text{g Li}}} = 8.07 \text{ mol Li}$$

$$56.0 \cancel{\text{g } N_2} \times \frac{1 \text{ mol } N_2}{28.02 \cancel{\text{g } N_2}} = 2.00 \text{ mol } N_2$$

Step 3 Using the mole ratio from the balanced equation, we can calculate the moles of lithium required to react with 2.00 moles of nitrogen.

$$2.00 \cancel{\text{mol } N_2} \times \frac{6 \text{ mol Li}}{1 \cancel{\text{mol } N_2}} = 12.0 \text{ mol Li}$$

Therefore, 12.0 moles of Li is required to react with 2.00 moles of N_2. However, we have only 8.07 mol of Li, so lithium is limiting. It will be consumed before the nitrogen runs out.

Step 4 Because lithium is the limiting reactant, we must use the 8.07 moles of Li to determine how many moles of Li_3N can be formed.

$$8.07 \cancel{\text{mol Li}} \times \frac{2 \text{ mol } Li_3N}{6 \cancel{\text{mol Li}}} = 2.69 \text{ mol } Li_3N$$

Step 5 We can now use the molar mass of Li_3N (34.83 g) to calculate the mass of Li_3N formed.

$$2.69 \cancel{\text{mol } Li_3N} \times \frac{34.83 \text{ g } Li_3N}{1 \cancel{\text{mol } Li_3N}} = 93.7 \text{ g } Li_3N$$

Self-Check Exercise 9.7

a. **Step 1** The balanced equation is

$$TiCl_4(g) + O_2(g) \rightarrow TiO_2(s) + 2Cl_2(g)$$

Step 2 The numbers of moles of reactants are

$$6.71 \times 10^3 \cancel{\text{g } TiCl_4} \times \frac{1 \text{ mol } TiCl_4}{189.68 \cancel{\text{g } TiCl_4}} = 3.54 \times 10^1 \text{ mol } TiCl_4$$

$$2.45 \times 10^3 \cancel{\text{g } O_2} \times \frac{1 \text{ mol } O_2}{32.00 \cancel{\text{g } O_2}} = 7.66 \times 10^1 \text{ mol } O_2$$

Step 3 In the balanced equation both $TiCl_4$ and O_2 have coefficients of 1, so

$$1 \text{ mol } TiCl_4 = 1 \text{ mol } O_2$$

and

$$3.54 \times 10^1 \cancel{\text{mol } TiCl_4} \times \frac{1 \text{ mol } O_2}{1 \cancel{\text{mol } TiCl_4}} = 3.54 \times 10^1 \text{ mol } O_2 \text{ required}$$

We have 7.66×10^1 moles of O_2, so the O_2 is in excess and the $TiCl_4$ is limiting. This makes sense. $TiCl_4$ and O_2 react in a 1:1 mole ratio, so the $TiCl_4$ is limiting because fewer moles of $TiCl_4$ are present than moles of O_2.

Step 4 We will now use the moles of $TiCl_4$ (the limiting reactant) to determine the moles of TiO_2 that would form if the reaction produced 100% of the expected yield (the theoretical yield).

$$3.54 \times 10^1 \cancel{\text{mol } TiCl_4} \times \frac{1 \text{ mol } TiO_2}{1 \cancel{\text{mol } TiCl_4}} = 3.54 \times 10^1 \text{ mol } TiO_2$$

The mass of TiO_2 expected for 100% yield is

$$3.54 \times 10^1 \cancel{\text{mol } TiO_2} \times \frac{79.88 \text{ g } TiO_2}{1 \cancel{\text{mol } TiO_2}} = 2.83 \times 10^3 \text{ g } TiO_2$$

This amount represents the theoretical yield.

b. Because the reaction is said to give only a 75.0% yield of TiO_2, we use the definition of percent yield,

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \%\text{ yield}$$

to write the equation

$$\frac{\text{Actual yield}}{2.83 \times 10^3 \text{ g TiO}_2} \times 100\% = 75.0\%\text{ yield}$$

We now want to solve for the actual yield. First we divide both sides by 100%.

$$\frac{\text{Actual yield}}{2.83 \times 10^3 \text{ g TiO}_2} \times \frac{100\%}{100\%} = \frac{75.0}{100} = 0.750$$

Then we multiply both sides by 2.83×10^3 g TiO_2.

$$2.83 \times 10^3 \text{ g TiO}_2 \times \frac{\text{Actual yield}}{2.83 \times 10^3 \text{ g TiO}_2} = 0.750 \times 2.83 \times 10^3 \text{ g TiO}_2$$

$$\text{Actual yield} = 0.750 \times 2.83 \times 10^3 \text{ g TiO}_2 = 2.12 \times 10^3 \text{ g TiO}_2$$

Thus 2.12×10^3 g of $TiO_2(s)$ is actually obtained in this reaction.

제 10 장

Self-Check Exercise 10.1

The conversion factor needed is $\frac{1 \text{ cal}}{4.184 \text{ J}}$, and the conversion is

$$28.4 \cancel{\text{J}} \times \frac{1 \text{ cal}}{4.184 \cancel{\text{J}}} = 6.79 \text{ cal}$$

Self-Check Exercise 10.2

We know that it takes 4.184 J of energy to change the temperature of each gram of water by 1 °C, so we must multiply 4.184 by the mass of water (454 g) and the temperature change (98.6 °C − 5.4 °C = 93.2 °C).

$$4.184 \frac{\text{J}}{\cancel{\text{g}}\ \cancel{°\text{C}}} \times 454 \cancel{\text{g}} \times 93.2 \cancel{°\text{C}} = 1.77 \times 10^5 \text{ J}$$

Self-Check Exercise 10.3

From Table 10.1, the specific heat capacity for solid gold is 0.13 J/g °C. Because it takes 0.13 J to change the temperature of *one* gram of gold by *one* Celsius degree, we must multiply 0.13 by the sample size (5.63 g) and the change in temperature (32 °C − 21 °C = 11 °C).

$$0.13 \frac{\text{J}}{\text{g °C}} \times 5.63 \text{ g} \times 11 \text{ °C} = 8.1 \text{ J}$$

We can change this energy to units of calories as follows:

$$8.1 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 1.9 \text{ cal}$$

Self-Check Exercise 10.4

Table 10.1 lists the specific heat capacities of several metals. We want to calculate the specific heat capacity (s) for this metal and then use Table 10.1 to identify the metal. Using the equation

$$Q = s \times m \times \Delta T$$

we can solve for s by dividing both sides by m (the mass of the sample) and by ΔT:

$$\frac{Q}{m \times \Delta T} = s$$

In this case,

Q = energy (heat) required = 10.1 J
m = 2.8 g
ΔT = temperature change = 36 °C − 21 °C = 15 °C

so

$$s = \frac{Q}{m \times \Delta T} = \frac{10.1 \text{ J}}{(2.8 \text{ g})(15 \text{ °C})} = 0.24 \text{ J/g °C}$$

Table 10.1 shows that silver has a specific heat capacity of 0.24 J/g °C. The metal is silver.

Self-Check Exercise 10.5

We are told that 1652 kJ of energy is *released* when 4 moles of Fe reacts. We first need to determine what number of moles 1.00 g Fe represents.

$$1.00 \text{ g Fe} \times \frac{1 \text{ mol}}{55.85 \text{ g}} = 1.79 \times 10^{-2} \text{ mol Fe}$$

$$1.79 \times 10^{-2} \text{ mol Fe} \times \frac{1652 \text{ kJ}}{4 \text{ mol Fe}} = 7.39 \text{ kJ}$$

Thus 7.39 kJ of energy (as heat) is released when 1.00 g of iron reacts.

Self-Check Exercise 10.6

Noting the reactants and products in the desired reaction

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

We need to reverse the second equation and multiply it by $\frac{1}{2}$. This reverses the sign and cuts the amount of energy by a factor of 2.

$$\tfrac{1}{2}[2SO_3(g) \rightarrow 2SO_2(g) + O_2(g)] \qquad \Delta H = \frac{198.2 \text{ kJ}}{2}$$

or

$$SO_3(g) \rightarrow SO_2(g) + \tfrac{1}{2}O_2(g) \qquad \Delta H = 99.1 \text{ kJ}$$

Now we add this reaction to the first reaction.

$$\begin{array}{ll} S(s) + \tfrac{3}{2}O_2(g) \rightarrow SO_3(g) & \Delta H = -395.2 \text{ kJ} \\ SO_3(g) \rightarrow SO_2(g) + \tfrac{1}{2}O_2(g) & \Delta H = 99.1 \text{ kJ} \\ \hline S(s) + O_2(g) \rightarrow SO_2(g) & \Delta H = -296.1 \text{ kJ} \end{array}$$

제 11 장

Self-Check Exercise 11.1

Using the electronegativity values given in Figure 11.3, we choose the bond in which the atoms exhibit the largest difference in electronegativity. (Electronegativity values are shown in parentheses.)

a. H—C (2.1)(2.5) > H—P (2.1)(2.1)

b. O—I (3.5)(2.5) > O—F (3.5)(4.0)

c. S—O (2.5)(3.5) > N—O (3.0)(3.5)

d. N—H (3.0)(2.1) > Si—H (1.8)(2.1)

Self-Check Exercise 11.2

H has one electron, and Cl has seven valence electrons. This gives a total of eight valence electrons. We first draw in the bonding pair:

$H—Cl$, which could be drawn as $H:Cl$

We have six electrons yet to place. The H already has two electrons, so we place three lone pairs around the chlorine to satisfy the octet rule.

$$H—\ddot{\underset{\cdot\cdot}{Cl}}: \text{ or } H:\ddot{\underset{\cdot\cdot}{Cl}}:$$

Self-Check Exercise 11.3

Step 1 O_3: 3(6) = 18 valence electrons
Step 2 O—O—O
Step 3 $\ddot{O}=\ddot{O}—\ddot{\underset{\cdot\cdot}{O}}:$ and $:\ddot{\underset{\cdot\cdot}{O}}—\ddot{O}=\ddot{O}$

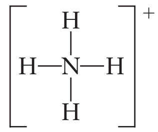
This molecule shows resonance (it has two valid Lewis structures).

Self-Check Exercise 11.4

See table on top of page A10.

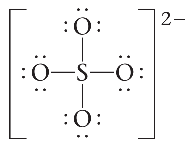
Self-Check Exercise 11.5

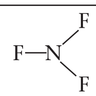
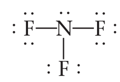
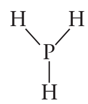
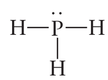
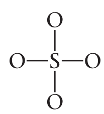
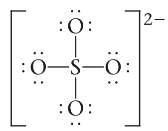
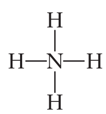
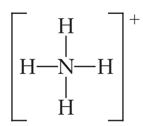
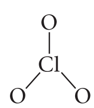
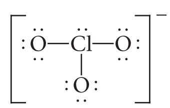
a. NH_4^+

The Lewis structure is 

(See Self-Check Exercise 11.4.) There are four pairs of electrons around the nitrogen. This requires a tetrahedral arrangement of electron pairs. The NH_4^+ ion has a tetrahedral molecular structure (row 3 in Table 11.4), because all electron pairs are shared.

b. SO_4^{2-}

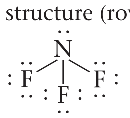
The Lewis structure is 

Molecule or Ion	Total Valence Electrons	Draw Single Bonds	Calculate Number of Electrons Remaining	Use Remaining Electrons to Achieve Noble Gas Configurations	Check: Atoms	Check: Electrons
a. NF_3	5 + 3(7) = 26		26 − 6 = 20		N F	8 8
b. O_2	2(6) = 12	O—O	12 − 2 = 10	$\dot{\underset{\cdot}{\ddot{O}}}=\dot{\underset{\cdot}{\ddot{O}}}$	O	8
c. CO	4 + 6 = 10	C—O	10 − 2 = 8	:C≡O:	C O	8 8
d. PH_3	5 + 3(1) = 8		8 − 6 = 2		P H	8 2
e. H_2S	2(1) + 6 = 8	H—S—H	8 − 4 = 4	$H—\ddot{\underset{\cdot\cdot}{S}}—H$	S H	8 2
f. SO_4^{2-}	6 + 4(6) + 2 = 32		32 − 8 = 24		S O	8 8
g. NH_4^+	5 + 4(1) − 1 = 8		8 − 8 = 0		N H	8 2
h. ClO_3^-	7 + 3(6) + 1 = 26		26 − 6 = 20		Cl O	8 8
i. SO_2	6 + 2(6) = 18	O—S—O	18 − 4 = 14	$\ddot{O}=\ddot{S}—\ddot{\underset{\cdot\cdot}{O}}:$ and $:\ddot{\underset{\cdot\cdot}{O}}—\ddot{S}=\ddot{O}$	S O	8 8

Answer to Self-Check Exercise 11.4.

(See Self-Check Exercise 11.4.) The four electron pairs around the sulfur require a tetrahedral arrangement. The SO_4^{2-} has a tetrahedral molecular structure (row 3 in Table 11.4).

c. NF_3

The Lewis structure is 

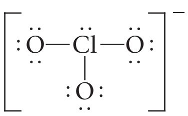
(See Self-Check Exercise 11.4.) The four pairs of electrons on the nitrogen require a tetrahedral arrangement. In this case, only three of the pairs are shared with the fluorine atoms, leaving one lone pair. Thus the molecular structure is a trigonal pyramid (row 4 in Table 11.4).

d. H_2S

The Lewis structure is H—S̈—H

(See Self-Check Exercise 11.4.) The four pairs of electrons around the sulfur require a tetrahedral arrangement. In this case, two pairs are shared with hydrogen atoms, leaving two lone pairs. Thus the molecular structure is bent or V-shaped (row 5 in Table 11.4).

e. ClO_3^-

The Lewis structure is 

(See Self-Check Exercise 11.4.) The four pairs of electrons require a tetrahedral arrangement. In this case, three pairs are shared with oxygen atoms, leaving one lone pair. Thus the molecular structure is a trigonal pyramid (row 4 in Table 11.4).

f. BeF_2

The Lewis structure is :F̈—Be—F̈:

The two electron pairs on beryllium require a linear arrangement. Because both pairs are shared by fluorine atoms, the molecular structure is also linear (row 1 in Table 11.4).

제 12 장

Self-Check Exercise 12.1

We know that 1.000 atm = 760.0 mm Hg. So

$$525\ \cancel{\text{mm Hg}} \times \frac{1.000\ \text{atm}}{760.0\ \cancel{\text{mm Hg}}} = 0.691\ \text{atm}$$

Self-Check Exercise 12.2

Initial Conditions	**Final Conditions**
$P_1 = 635$ torr	$P_2 = 785$ torr
$V_1 = 1.51$ L	$V_2 = ?$

Solving Boyle's law ($P_1V_1 = P_2V_2$) for V_2 gives

$$V_2 = V_1 \times \frac{P_1}{P_2}$$
$$= 1.51\ \text{L} \times \frac{635\ \cancel{\text{torr}}}{785\ \cancel{\text{torr}}} = 1.22\ \text{L}$$

Note that the volume decreased, as the increase in pressure led us to expect.

Self-Check Exercise 12.3

Because the temperature of the gas inside the bubble decreases (at constant pressure), the bubble gets smaller. The conditions are

Initial Conditions

$T_1 = 28\ °C = 28 + 273 = 301$ K

$V_1 = 23\ \text{cm}^3$

Final Conditions

$T_2 = 18\ °C = 18 + 273 = 291$ K

$V_2 = ?$

Solving Charles's law,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

for V_2 gives

$$V_2 = V_1 \times \frac{T_2}{T_1} = 23\ \text{cm}^3 \times \frac{291\ \cancel{\text{K}}}{301\ \cancel{\text{K}}} = 22\ \text{cm}^3$$

Self-Check Exercise 12.4

Because the temperature and pressure of the two samples are the same, we can use Avogadro's law in the form

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

The following information is given:

Sample 1	**Sample 2**
$V_1 = 36.7$ L	$V_2 = 16.5$ L
$n_1 = 1.5$ mol	$n_2 = ?$

We can now solve Avogadro's law for the value of n_2 (the moles of N_2 in sample 2):

$$n_2 = n_1 \times \frac{V_2}{V_1} = 1.5\ \text{mol} \times \frac{16.5\ \cancel{\text{L}}}{36.7\ \cancel{\text{L}}} = 0.67\ \text{mol}$$

Here n_2 is smaller than n_1, which makes sense in view of the fact that V_2 is smaller than V_1.

Note: We isolate n_2 from Avogadro's law as given above by multiplying both sides of the equation by n_2 and then by n_1/V_1,

$$\left(n_2 \times \frac{n_1}{V_1}\right)\frac{V_1}{n_1} = \left(n_2 \times \frac{n_1}{V_1}\right)\frac{V_2}{n_2}$$

to give $n_2 = n_1 \times V_2/V_1$.

Self-Check Exercise 12.5

We are given the following information:

$P = 1.00$ atm

$V = 2.70 \times 10^6$ L

$n = 1.10 \times 10^5$ mol

We solve for T by dividing both sides of the ideal gas law by nR:

$$\frac{PV}{nR} = \frac{\cancel{nR}T}{\cancel{nR}}$$

to give

$$T = \frac{PV}{nR} = \frac{(1.00\ \cancel{\text{atm}})(2.70 \times 10^6\ \cancel{\text{L}})}{(1.10 \times 10^5\ \cancel{\text{mol}})\left(0.08206\ \dfrac{\cancel{\text{L}}\ \cancel{\text{atm}}}{\text{K}\ \cancel{\text{mol}}}\right)}$$
$$= 299\ \text{K}$$

The temperature of the helium is 299 K, or 299 − 273 = 26 °C.

Self-Check Exercise 12.6

We are given the following information about the radon sample:

$n = 1.5$ mol

$V = 21.0$ L

$T = 33\ °C = 33 + 273 = 306$ K

$P = ?$

We solve the ideal gas law ($PV = nRT$) for P by dividing both sides of the equation by V:

$$P = \frac{nRT}{V} = \frac{(1.5\ \cancel{\text{mol}})\left(0.08206\ \frac{\cancel{\text{L}}\ \text{atm}}{\cancel{\text{K}}\ \cancel{\text{mol}}}\right)(306\ \cancel{\text{K}})}{21.0\ \cancel{\text{L}}}$$
$$= 1.8\ \text{atm}$$

Self-Check Exercise 12.7

We are given the following information:

Initial Conditions

$P_1 = 0.747\ \text{atm}$
$T_1 = 13\ °\text{C} = 13 + 273 = 286\ \text{K}$
$V_1 = 11.0\ \text{L}$

Final Conditions

$P_2 = 1.18\ \text{atm}$
$T_2 = 56\ °\text{C} = 56 + 273 = 329\ \text{K}$
$V_2 = ?$

In this case, the number of moles remains constant. Thus we can say

$$\frac{P_1V_1}{T_1} = nR \quad \text{and} \quad \frac{P_2V_2}{T_2} = nR$$

or

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Solving for V_2 gives

$$V_2 = V_1 \times \frac{T_2}{T_1} \times \frac{P_1}{P_2} = (11.0\ \text{L})\left(\frac{329\ \cancel{\text{K}}}{286\ \cancel{\text{K}}}\right)\left(\frac{0.747\ \cancel{\text{atm}}}{1.18\ \cancel{\text{atm}}}\right)$$
$$= 8.01\ \text{L}$$

Self-Check Exercise 12.8

As usual when dealing with gases, we can use the ideal gas equation $PV = nRT$. First consider the information given:

$P = 0.91\ \text{atm} = P_{\text{total}}$
$V = 2.0\ \text{L}$
$T = 25\ °\text{C} = 25 + 273 = 298\ \text{K}$

Given this information, we can calculate the number of moles of gas in the mixture: $n_{\text{total}} = n_{N_2} + n_{O_2}$. Solving for n in the ideal gas equation gives

$$n_{\text{total}} = \frac{P_{\text{total}}V}{RT} = \frac{(0.91\ \cancel{\text{atm}})(2.0\ \cancel{\text{L}})}{\left(0.08206\ \frac{\cancel{\text{L}}\ \cancel{\text{atm}}}{\cancel{\text{K}}\ \text{mol}}\right)(298\ \cancel{\text{K}})} = 0.074\ \text{mol}$$

We also know that 0.050 mole of N_2 is present. Because

$$n_{\text{total}} = \underset{\substack{\uparrow \\ (0.050\ \text{mol})}}{n_{N_2}} + n_{O_2} = 0.074\ \text{mol}$$

we can calculate the moles of O_2 present.

$$0.050\ \text{mol} + n_{O_2} = 0.074\ \text{mol}$$
$$n_{O_2} = 0.074\ \text{mol} - 0.050\ \text{mol} = 0.024\ \text{mol}$$

Now that we know the moles of oxygen present, we can calculate the partial pressure of oxygen from the ideal gas equation.

$$P_{O_2} = \frac{n_{O_2}RT}{V} = \frac{(0.024\ \cancel{\text{mol}})\left(0.08206\ \frac{\cancel{\text{L}}\ \text{atm}}{\cancel{\text{K}}\ \cancel{\text{mol}}}\right)(298\ \cancel{\text{K}})}{2.0\ \cancel{\text{L}}}$$
$$= 0.29\ \text{atm}$$

Although it is not requested, note that the partial pressure of the N_2 must be 0.62 atm, because

$$\underbrace{0.62\ \text{atm}}_{P_{N_2}} + \underbrace{0.29\ \text{atm}}_{P_{O_2}} = \underbrace{0.91\ \text{atm}}_{P_{\text{total}}}$$

Self-Check Exercise 12.9

The volume is 0.500 L, the temperature is 25 °C (or 25 + 273 = 298 K), and the total pressure is given as 0.950 atm. Of this total pressure, 24 torr is due to the water vapor. We can calculate the partial pressure of the H_2 because we know that

$$P_{\text{total}} = P_{H_2} + \underset{\substack{\uparrow \\ \text{24 torr}}}{P_{H_2O}} = 0.950\ \text{atm}$$

Before we carry out the calculation, however, we must convert the pressures to the same units. Converting P_{H_2O} to atmospheres gives

$$24\ \cancel{\text{torr}} \times \frac{1.000\ \text{atm}}{760.0\ \cancel{\text{torr}}} = 0.032\ \text{atm}$$

Thus

$$P_{\text{total}} = P_{H_2} + P_{H_2O} = 0.950\ \text{atm} = P_{H_2} + 0.032\ \text{atm}$$

and

$$P_{H_2} = 0.950\ \text{atm} - 0.032\ \text{atm} = 0.918\ \text{atm}$$

Now that we know the partial pressure of the hydrogen gas, we can use the ideal gas equation to calculate the moles of H_2.

$$n_{H_2} = \frac{P_{H_2}V}{RT} = \frac{(0.918\ \cancel{\text{atm}})(0.500\ \cancel{\text{L}})}{\left(0.08206\ \frac{\cancel{\text{L}}\ \cancel{\text{atm}}}{\cancel{\text{K}}\ \text{mol}}\right)(298\ \cancel{\text{K}})}$$
$$= 0.0188\ \text{mol} = 1.88 \times 10^{-2}\ \text{mol}$$

The sample of gas contains 1.88×10^{-2} mole of H_2, which exerts a partial pressure of 0.918 atm.

Self-Check Exercise 12.10

We will solve this problem by taking the following steps:

Grams of zinc ⇨ Moles of zinc ⇨ Moles of H_2 ⇨ Volume of H_2

Step 1 Using the atomic mass of zinc (65.38), we calculate the moles of zinc in 26.5 g.

$$26.5\ \cancel{\text{g Zn}} \times \frac{1\ \text{mol Zn}}{65.38\ \cancel{\text{g Zn}}} = 0.405\ \text{mol Zn}$$

Step 2 Using the balanced equation, we next calculate the moles of H_2 produced.

$$0.405\ \cancel{\text{mol Zn}} \times \frac{1\ \text{mol}\ H_2}{1\ \cancel{\text{mol Zn}}} = 0.405\ \text{mol}\ H_2$$

Step 3 Now that we know the moles of H_2, we can compute the volume of H_2 by using the ideal gas law, where

$P = 1.50\ \text{atm}$
$V = ?$
$n = 0.405\ \text{mol}$
$R = 0.08206\ \text{L atm/K mol}$
$T = 19\ °\text{C} = 19 + 273 = 292\ \text{K}$

$$V = \frac{nRT}{P} = \frac{(0.405\ \cancel{\text{mol}})\left(0.08206\ \frac{\text{L}\ \cancel{\text{atm}}}{\cancel{\text{K}}\ \cancel{\text{mol}}}\right)(292\ \cancel{\text{K}})}{1.50\ \cancel{\text{atm}}}$$
$$= 6.47\ \text{L of}\ H_2$$

Self-Check Exercise 12.11

Although there are several possible ways to do this problem, the most convenient method involves using the molar volume at STP. First we use the ideal gas equation to calculate the moles of NH_3 present:

$$n = \frac{PV}{RT}$$

where $P = 15.0$ atm, $V = 5.00$ L, and $T = 25\ °C + 273 = 298$ K.

$$n = \frac{(15.0\ \cancel{\text{atm}})(5.00\ \cancel{\text{L}})}{\left(0.08206\ \frac{\cancel{\text{L}}\ \cancel{\text{atm}}}{\cancel{\text{K}}\ \text{mol}}\right)(298\ \cancel{\text{K}})} = 3.07\ \text{mol}$$

We know that at STP each mole of gas occupies 22.4 L. Therefore, 3.07 mol has the volume

$$3.07\ \cancel{\text{mol}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 68.8\ \text{L}$$

The volume of the ammonia at STP is 68.8 L.

제 13 장

Self-Check Exercise 13.1

Energy to melt the ice:

$$15\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \text{mol H}_2\text{O}}{18\ \cancel{\text{g H}_2\text{O}}} = 0.83\ \text{mol H}_2\text{O}$$

$$0.83\ \cancel{\text{mol H}_2\text{O}} \times 6.02\ \frac{\text{kJ}}{\cancel{\text{mol H}_2\text{O}}} = 5.0\ \text{kJ}$$

Energy to heat the water from 0 °C to 100 °C:

$$4.18\ \frac{\text{J}}{\cancel{\text{g}}\ \cancel{°\text{C}}} \times 15\ \cancel{\text{g}} \times 100\ \cancel{°\text{C}} = 6300\ \text{J}$$

$$6300\ \cancel{\text{J}} \times \frac{1\ \text{kJ}}{1000\ \cancel{\text{J}}} = 6.3\ \text{kJ}$$

Energy to vaporize the water at 100 °C:

$$0.83\ \cancel{\text{mol H}_2\text{O}} \times 40.6\ \frac{\text{kJ}}{\cancel{\text{mol H}_2\text{O}}} = 34\ \text{kJ}$$

Total energy required:

$$5.0\ \text{kJ} + 6.3\ \text{kJ} + 34\ \text{kJ} = 45\ \text{kJ}$$

Self-Check Exercise 13.2

a. Contains SO_3 molecules—a molecular solid.
b. Contains Ba^{2+} and O^{2-} ions—an ionic solid.
c. Contains Au atoms—an atomic solid.

제 14 장

Self-Check Exercise 14.1

$$\text{Mass percent} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

For this sample, the mass of solution is 135 g and the mass of the solute is 4.73 g, so

$$\text{Mass percent} = \frac{4.73\ \text{g solute}}{135\ \text{g solution}} \times 100\% = 3.50\%$$

Self-Check Exercise 14.2

Using the definition of mass percent, we have

$$\frac{\text{Mass of solute}}{\text{Mass of solution}} = \frac{\text{grams of solute}}{\text{grams of solute} + \text{grams of solvent}} \times 100\% = 40.0\%$$

There are 425 grams of solute (formaldehyde). Substituting, we have

$$\frac{425\ \text{g}}{425\ \text{g} + \text{grams of solvent}} \times 100\% = 40.0\%$$

We must now solve for grams of solvent (water). This will take some patience, but we can do it if we proceed step by step. First we divide both sides by 100%.

$$\frac{425\ \text{g}}{425\ \text{g} + \text{grams of solvent}} \times \frac{\cancel{100\%}}{\cancel{100\%}} = \frac{40.0\%}{100\%} = 0.400$$

Now we have

$$\frac{425\ \text{g}}{425\ \text{g} + \text{grams of solvent}} = 0.400$$

Next we multiply both sides by (425 g + grams of solvent).

$$\cancel{(425\ \text{g} + \text{grams of solvent})} \times \frac{425\ \text{g}}{\cancel{425\ \text{g} + \text{grams of solvent}}} = 0.400 \times (425\ \text{g} + \text{grams of solvent})$$

This gives

$$425\ \text{g} = 0.400 \times (425\ \text{g} + \text{grams of solvent})$$

Carrying out the multiplication gives

$$425\ \text{g} = 170.\ \text{g} + 0.400\ (\text{grams of solvent})$$

Now we subtract 170. g from both sides,

$$425\ \text{g} - 170.\ \text{g} = \cancel{170.\ \text{g}} - \cancel{170.\ \text{g}} + 0.400\ (\text{grams of solvent})$$

$$255\ \text{g} = 0.400\ (\text{grams of solvent})$$

and divide both sides by 0.400.

$$\frac{255\ \text{g}}{0.400} = \frac{\cancel{0.400}}{\cancel{0.400}}\ (\text{grams of solvent})$$

We finally have the answer:

$$\frac{255\ \text{g}}{0.400} = 638\ \text{g} = \text{grams of solvent} = \text{mass of water needed}$$

Self-Check Exercise 14.3

The moles of ethanol can be obtained from its molar mass (46.1).

$$1.00\ \cancel{\text{g C}_2\text{H}_5\text{OH}} \times \frac{1\ \text{mol C}_2\text{H}_5\text{OH}}{46.1\ \cancel{\text{g C}_2\text{H}_5\text{OH}}} = 2.17 \times 10^{-2}\ \text{mol C}_2\text{H}_5\text{OH}$$

$$\text{Volume in liters} = 101\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.101\ \text{L}$$

$$\text{Molarity of C}_2\text{H}_5\text{OH} = \frac{\text{moles of C}_2\text{H}_5\text{OH}}{\text{liters of solution}} = \frac{2.17 \times 10^{-2}\ \text{mol}}{0.101\ \text{L}} = 0.215\ M$$

Self-Check Exercise 14.4

When Na_2CO_3 and $Al_2(SO_4)_3$ dissolve in water, they produce ions as follows:

$$\text{Na}_2\text{CO}_3(s) \xrightarrow{\text{H}_2\text{O}(l)} 2\text{Na}^+(aq) + \text{CO}_3^{2-}(aq)$$

$$\text{Al}_2(\text{SO}_4)_3(s) \xrightarrow{\text{H}_2\text{O}(l)} 2\text{Al}^{3+}(aq) + 3\text{SO}_4^{2-}(aq)$$

Therefore, in a 0.10 M Na_2CO_3 solution, the concentration of Na^+ ions is $2 \times 0.10\ M = 0.20\ M$ and the concentration of CO_3^{2-} ions is 0.10 M. In a 0.010 M $Al_2(SO_4)_3$ solution, the concentration of Al^{3+} ions is $2 \times 0.010\ M = 0.020\ M$ and the concentration of SO_4^{2-} ions is $3 \times 0.010\ M = 0.030\ M$.

Self-Check Exercise 14.5

When solid $AlCl_3$ dissolves, it produces ions as follows:

$$\text{AlCl}_3(s) \xrightarrow{\text{H}_2\text{O}(l)} \text{Al}^{3+}(aq) + 3\text{Cl}^-(aq)$$

so a 1.0×10^{-3} M $AlCl_3$ solution contains 1.0×10^{-3} M Al^{3+} ions and 3.0×10^{-3} M Cl^- ions.

To calculate the moles of Cl^- ions in 1.75 L of the 1.0×10^{-3} M $AlCl_3$ solution, we must multiply the volume by the molarity.

$$1.75\ \text{L solution} \times 3.0 \times 10^{-3}\ M\ \text{Cl}^-$$

$$= 1.75\ \cancel{\text{L solution}} \times \frac{3.0 \times 10^{-3}\ \text{mol Cl}^-}{\cancel{\text{L solution}}}$$
$$= 5.25 \times 10^{-3}\ \text{mol Cl}^- = 5.3 \times 10^{-3}\ \text{mol Cl}^-$$

Self-Check Exercise 14.6

We must first determine the number of moles of formaldehyde in 2.5 L of 12.3 M formalin. Remember that volume of solution (in liters) times molarity gives moles of solute. In this case, the volume of solution is 2.5 L and the molarity is 12.3 moles of HCHO per liter of solution.

$$2.5\ \cancel{\text{L solution}} \times \frac{12.3\ \text{mol HCHO}}{\cancel{\text{L solution}}} = 31\ \text{mol HCHO}$$

Next, using the molar mass of HCHO (30.0 g), we convert 31 moles of HCHO to grams.

$$31\ \cancel{\text{mol HCHO}} \times \frac{30.0\ \text{g HCHO}}{1\ \cancel{\text{mol HCHO}}} = 9.3 \times 10^2\ \text{g HCHO}$$

Therefore, 2.5 L of 12.3 M formalin contains 9.3×10^2 g of formaldehyde. We must weigh out 930 g of formaldehyde and dissolve it in enough water to make 2.5 L of solution.

Self-Check Exercise 14.7

We are given the following information:

$M_1 = 12\,\frac{\text{mol}}{\text{L}}$ $\quad M_2 = 0.25\,\frac{\text{mol}}{\text{L}}$

$V_1 = ?$ (what we need to find) $\quad V_2 = 0.75$ L

Using the fact that the moles of solute do not change upon dilution, we know that

$$M_1 \times V_1 = M_2 \times V_2$$

Solving for V_1 by dividing both sides by M_1 gives

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.25\,\frac{\cancel{\text{mol}}}{\cancel{\text{L}}} \times 0.75\ \text{L}}{12\,\frac{\cancel{\text{mol}}}{\cancel{\text{L}}}}$$

and

$$V_1 = 0.016\ \text{L} = 16\ \text{mL}$$

Self-Check Exercise 14.8

Step 1 When the aqueous solutions of Na_2SO_4 (containing Na^+ and SO_4^{2-} ions) and $Pb(NO_3)_2$ (containing Pb^{2+} and NO_3^- ions) are mixed, solid $PbSO_4$ is formed.

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

Step 2 We must first determine whether Pb^{2+} or SO_4^{2-} is the limiting reactant by calculating the moles of Pb^{2+} and SO_4^{2-} ions present. Because 0.0500 M $Pb(NO_3)_2$ contains 0.0500 M Pb^{2+} ions, we can calculate the moles of Pb^{2+} ions in 1.25 L of this solution as follows:

$$1.25\ \cancel{\text{L}} \times \frac{0.0500\ \text{mol Pb}^{2+}}{\cancel{\text{L}}} = 0.0625\ \text{mol Pb}^{2+}$$

The 0.0250 M Na_2SO_4 solution contains 0.0250 M SO_4^{2-} ions, and the number of moles of SO_4^{2-} ions in 2.00 L of this solution is

$$2.00\ \cancel{\text{L}} \times \frac{0.0250\ \text{mol SO}_4^{2-}}{\cancel{\text{L}}} = 0.0500\ \text{mol SO}_4^{2-}$$

Step 3 Pb^{2+} and SO_4^{2-} react in a 1:1 ratio, so the amount of SO_4^{2-} ions is limiting because SO_4^{2-} is present in the smaller number of moles.

Step 4 The Pb^{2+} ions are present in excess, and only 0.0500 mole of solid $PbSO_4$ will be formed.

Step 5 We calculate the mass of $PbSO_4$ by using the molar mass of $PbSO_4$ (303.3 g).

$$0.0500\ \cancel{\text{mol PbSO}_4} \times \frac{303.3\ \text{g PbSO}_4}{1\ \cancel{\text{mol PbSO}_4}} = 15.2\ \text{g PbSO}_4$$

Self-Check Exercise 14.9

Step 1 Because nitric acid is a strong acid, the nitric acid solution contains H^+ and NO_3^- ions. The KOH solution contains K^+ and OH^- ions. When these solutions are mixed, the H^+ and OH^- react to form water.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Step 2 The number of moles of OH^- present in 125 mL of 0.050 M KOH is

$$125\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.050\ \text{mol OH}^-}{\cancel{\text{L}}} = 6.3 \times 10^{-3}\ \text{mol OH}^-$$

Step 3 H^+ and OH^- react in a 1:1 ratio, so we need 6.3×10^{-3} mole of H^+ from the 0.100 M HNO_3.

Step 4 6.3×10^{-3} mole of OH^- requires 6.3×10^{-3} mole of H^+ to form 6.3×10^{-3} mole of H_2O.

Therefore,

$$V \times \frac{0.100\ \text{mol H}^+}{\text{L}} = 6.3 \times 10^{-3}\ \text{mol H}^+$$

where V represents the volume in liters of 0.100 M HNO_3 required. Solving for V, we have

$$V = \frac{6.3 \times 10^{-3}\ \cancel{\text{mol H}^+}}{\frac{0.100\ \cancel{\text{mol H}^+}}{\text{L}}} = 6.3 \times 10^{-2}\ \text{L}$$
$$= 6.3 \times 10^{-2}\ \cancel{\text{L}} \times \frac{1000\ \text{mL}}{\cancel{\text{L}}} = 63\ \text{mL}$$

Self-Check Exercise 14.10

From the definition of normality, N = equiv/L, we need to calculate (1) the equivalents of KOH and (2) the volume of the solution in liters. To find the number of equivalents, we use the equivalent weight of KOH, which is 56.1 g (see Table 15.2).

$$23.6\ \cancel{\text{g KOH}} \times \frac{1\ \text{equiv KOH}}{56.1\ \cancel{\text{g KOH}}} = 0.421\ \text{equiv KOH}$$

Next we convert the volume to liters.

$$755\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.755\ \text{L}$$

Finally, we substitute these values into the equation that defines normality.

$$\text{Normality} = \frac{\text{equiv}}{\text{L}} = \frac{0.421\ \text{equiv}}{0.755\ \text{L}} = 0.558\ N$$

Self-Check Exercise 14.11

To solve this problem, we use the relationship

$$N_{acid} \times V_{acid} = N_{base} \times V_{base}$$

where

$$N_{acid} = 0.50\,\frac{\text{equiv}}{\text{L}}$$
$$V_{acid} = ?$$
$$N_{base} = 0.80\,\frac{\text{equiv}}{\text{L}}$$
$$V_{base} = 0.250\ \text{L}$$

We solve the equation

$$N_{acid} \times V_{acid} = N_{base} \times V_{base}$$

for V_{acid} by dividing both sides by N_{acid}.

$$\frac{N_{acid} \times V_{acid}}{N_{acid}} = \frac{N_{base} \times V_{base}}{N_{acid}}$$

$$V_{acid} = \frac{N_{base} \times V_{base}}{N_{acid}} = \frac{\left(0.80\,\frac{\cancel{\text{equiv}}}{\cancel{\text{L}}}\right) \times (0.250\ \text{L})}{0.50\,\frac{\cancel{\text{equiv}}}{\cancel{\text{L}}}}$$

$$V_{acid} = 0.40\ \text{L}$$

Therefore, 0.40 L of 0.50 N H_2SO_4 is required to neutralize 0.250 L of 0.80 N KOH.

제 15 장

Self-Check Exercise 15.1

The conjugate acid–base pairs are

H_2O, H_3O^+
Base Conjugate acid

and

$HC_2H_3O_2$, $C_2H_3O_2^-$
Acid Conjugate base

The members of both pairs differ by one H^+.

Self-Check Exercise 15.2

Because $[H^+][OH^-] = 1.0 \times 10^{-14}$, we can solve for $[H^+]$.

$$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-2}} = 5.0 \times 10^{-13}\ M$$

This solution is basic: $[OH^-] = 2.0 \times 10^{-2}\ M$ is greater than $[H^+] = 5.0 \times 10^{-13}\ M$.

Self-Check Exercise 15.3

a. Because $[H^+] = 1.0 \times 10^{-3}\ M$, we get pH = 3.00 because $pH = -\log[H^+] = -\log[1.0 \times 10^{-3}] = 3.00$.

b. Because $[OH^-] = 5.0 \times 10^{-5}\ M$, we can find $[H^+]$ from the K_w expression.

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-5}} = 2.0 \times 10^{-10}\ M$$

$pH = -\log[H^+] = -\log[2.0 \times 10^{-10}] = 9.70$

Self-Check Exercise 15.4

$$pOH + pH = 14.00$$
$$pOH = 14.00 - pH = 14.00 - 3.5$$
$$pOH = 10.5$$

Self-Check Exercise 15.5

Step 1 pH = 3.50
Step 2 −pH = −3.50
Step 3 $\boxed{\text{inv}}$ $\boxed{\text{log}}$ $-3.50 = 3.2 \times 10^{-4}$
$[H^+] = 3.2 \times 10^{-4}\ M$

Self-Check Exercise 15.6

Step 1 pOH = 10.50
Step 2 −pOH = −10.50
Step 3 $\boxed{\text{inv}}$ $\boxed{\text{log}}$ $-10.50 = 3.2 \times 10^{-11}$
$[OH^-] = 3.2 \times 10^{-11}\ M$

Self-Check Exercise 15.7

Because HCl is a strong acid, it is completely dissociated:

$5.0 \times 10^{-3}\ M$ HCl $\rightarrow$ $5.0 \times 10^{-3}\ M$ H^+ and $5.0 \times 10^{-3}\ M$ Cl^-

so $[H^+] = 5.0 \times 10^{-3}\ M$.

$pH = -\log(5.0 \times 10^{-3}) = 2.30$

제 16 장

Self-Check Exercise 16.1

Applying the law of chemical equilibrium gives

Coefficient of NO_2 Coefficient of H_2O

$$K = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7}$$

Coefficient of O_2

Coefficient of NH_3

Self-Check Exercise 16.2

a. $K = [O_2]^3$ The solids are not included.

b. $K = [N_2O][H_2O]^2$ The solid is not included. Water is gaseous in this reaction, so it is included.

c. $K = \dfrac{1}{[CO_2]}$ The solids are not included.

d. $K = \dfrac{1}{[SO_3]}$ Water and H_2SO_4 are pure liquids and so are not included.

Self-Check Exercise 16.3

When rain is imminent, the concentration of water vapor in the air increases. This shifts the equilibrium to the right, forming $CoCl_2 \cdot 6H_2O(s)$, which is pink.

Self-Check Exercise 16.4

a. No change. Both sides of the equation contain the same number of gaseous components. The system cannot change its pressure by shifting its equilibrium position.

b. Shifts to the left. The system can increase the number of gaseous components present, and so increase the pressure, by shifting to the left.

c. Shifts to the right to increase the number of gaseous components and thus its pressure.

Self-Check Exercise 16.5

a. Shifts to the right away from added SO_2.

b. Shifts to the right to replace removed SO_3.

c. Shifts to the right to decrease its pressure.

d. Shifts to the right. Energy is a product in this case, so a decrease in temperature favors the forward reaction (which produces energy).

Self-Check Exercise 16.6

a. $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$; $K_{sp} = [Ba^{2+}][SO_4^{2-}]$

b. $Fe(OH)_3(s) \rightleftharpoons Fe^{3+}(aq) + 3OH^-(aq)$; $K_{sp} = [Fe^{3+}][OH^-]^3$

c. $Ag_3PO_4(s) \rightleftharpoons 3Ag^+(aq) + PO_4^{3-}(aq)$; $K_{sp} = [Ag^+]^3[PO_4^{3-}]$

Self-Check Exercise 16.7

$(3.9 \times 10^{-5})^2 = 1.5 \times 10^{-9} = K_{sp}$

Self-Check Exercise 16.8

$PbCrO_4(s) \rightleftharpoons Pb^{2+}(aq) + CrO_4^{2-}(aq)$
$K_{sp} = [Pb^{2+}][CrO_4^{2-}] = 2.0 \times 10^{-16}$
$[Pb^{2+}] = x$
$[CrO_4^{2-}] = x$
$K_{sp} = 2.0 \times 10^{-16} = x^2$
$x = [Pb^{2+}] = [CrO_4^{2-}] = 1.4 \times 10^{-8}$

제 17 장

Self-Check Exercise 17.1

a. CuO contains Cu^{2+} and O^{2-} ions, so copper is oxidized ($Cu \rightarrow Cu^{2+} + 2e^-$) and oxygen is reduced ($O + 2e^- \rightarrow O^{2-}$).

b. CsF contains Cs^+ and F^- ions. Thus cesium is oxidized ($Cs \rightarrow Cs^+ + e^-$) and fluorine is reduced ($F + e^- \rightarrow F^-$).

Self-Check Exercise 17.2

a. SO_3

We assign oxygen first. Each O is assigned an oxidation state of -2, giving a total of -6 (3×-2) for the three oxygen atoms. Because the molecule has zero charge overall, the sulfur must have an oxidation state of $+6$.

Check: $+6 + 3(-2) = 0$

b. SO_4^{2-}

As in part a, each oxygen is assigned an oxidation state of -2, giving a total of -8 (4×-2) on the four oxygen atoms. The anion has a net charge of -2, so the sulfur must have an oxidation state of $+6$.

Check: $+6 + 4(-2) = -2$

SO_4^{2-} has a charge of -2, so this is correct.

c. N_2O_5

We assign oxygen before nitrogen because oxygen is more electronegative. Thus each O is assigned an oxidation state of -2, giving a total of -10 (5×-2) on the five oxygen atoms. Therefore, the oxidation states of the *two* nitrogen atoms must total $+10$ because N_2O_5 has no overall charge. Each N is assigned an oxidation state of $+5$.

Check: $2(+5) + 5(-2) = 0$

d. PF_3

First we assign the fluorine an oxidation state of -1, giving a total of -3 (3×-1) on the three fluorine atoms. Thus P must have an oxidation state of $+3$.

Check: $+3 + 3(-1) = 0$

e. C_2H_6

In this case, it is best to recognize that hydrogen is always $+1$ in compounds with nonmetals. Thus each H is assigned an oxidation state of $+1$, which means that the six H atoms account for a total of $+6$ ($6 \times +1$). Therefore, the *two* carbon atoms must account for -6, and each carbon is assigned an oxidation state of -3.

Check: $2(-3) + 6(+1) = 0$

Self-Check Exercise 17.3

We can tell whether this is an oxidation–reduction reaction by comparing the oxidation states of the elements in the reactants and products:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Oxidation states: 0 0 −3 +1 (each H)

Nitrogen goes from 0 to -3. Thus it gains three electrons and is reduced. Each hydrogen atom goes from 0 to $+1$ and is thus oxidized, so this is an oxidation–reduction reaction. The oxidizing agent is N_2 (it takes electrons from H_2). The reducing agent is H_2 (it gives electrons to N_2).

$$N_2 + 3H_2 \rightarrow 2NH_3$$

$6e^-$

Self-Check Exercise 17.4

The unbalanced equation for this reaction is

$$Cu(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O(l) + NO(g)$$

Copper metal | Nitric acid | Aqueous copper(II) nitrate (contains Cu^{2+}) | Water | Nitrogen monoxide

Step 1 The oxidation half-reaction is

$$Cu + HNO_3 \rightarrow Cu(NO_3)_2$$

Oxidation states: 0 +1 +5 −2 (each O) +2 +5 −2 (each O)

The copper goes from 0 to $+2$ and thus is oxidized. This reduction reaction is

$$HNO_3 \rightarrow NO$$

Oxidation states: +1 +5 −2 +2 −2 (each O)

In this case, nitrogen goes from $+5$ in HNO_3 to $+2$ in NO and so is reduced. Notice two things about these reactions:

1. The HNO_3 must be included in the oxidation half-reaction to supply NO_3^- in the product $Cu(NO_3)_2$.
2. Although water is a product in the overall reaction, it does not need to be included in either half-reaction at the beginning. It will appear later as we balance the equation.

Step 2 *Balance the oxidation half-reaction.*

$$Cu + HNO_3 \rightarrow Cu(NO_3)_2$$

a. Balance nitrogen first.

$$Cu + 2HNO_3 \rightarrow Cu(NO_3)_2$$

b. Balancing nitrogen also caused oxygen to balance.

c. Balance hydrogen using H^+.

$$Cu + 2HNO_3 \rightarrow Cu(NO_3)_2 + 2H^+$$

d. Balance the charge using e^-.

$$Cu + 2HNO_3 \rightarrow Cu(NO_3)_2 + 2H^+ + 2e^-$$

This is the balanced oxidation half-reaction.

Balance the reduction half-reaction.

$$HNO_3 \rightarrow NO$$

a. All elements are balanced except hydrogen and oxygen.

b. Balance oxygen using H_2O.

$$HNO_3 \rightarrow NO + 2H_2O$$

c. Balance hydrogen using H^+.

$$3H^+ + HNO_3 \rightarrow NO + 2H_2O$$

d. Balance the charge using e^-.

$$3e^- + 3H^+ + HNO_3 \rightarrow NO + 2H_2O$$

This is the balanced reduction half-reaction.

Step 3 We equalize electrons by multiplying the oxidation half-reaction by 3:

$$3 \times [Cu + 2HNO_3 \rightarrow Cu(NO_3)_2 + 2H^+ + 2e^-]$$

gives

$$3Cu + 6HNO_3 \rightarrow 3Cu(NO_3)_2 + 6H^+ + 6e^-$$

Multiplying the reduction half-reaction by 2:

$$2 \times [3e^- + 3H^+ + HNO_3 \rightarrow NO + 2H_2O]$$

gives

$$6e^- + 6H^+ + 2HNO_3 \rightarrow 2NO + 4H_2O$$

Step 4 We can now add the balanced half-reactions, which both involve a six-electron change.

$$3Cu + 6HNO_3 \rightarrow 3Cu(NO_3)_2 + 6H^+ + 6e^-$$
$$6e^- + 6H^+ + 2HNO_3 \rightarrow 2NO + 4H_2O$$
$$\cancel{6e^-} + \cancel{6H^+} + 3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O + \cancel{6H^+} + \cancel{6e^-}$$

Canceling species common to both sides gives the balanced overall equation:

$$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$$

Step 5 Check the elements and charges.

Elements 3Cu, 8H, 8N, 24O $\rightarrow$ 3Cu, 8H, 8N, 24O
Charges 0 $\rightarrow$ 0

제 18 장

Self-Check Exercise 18.1

a. An alpha particle is a helium nucleus, ${}^{4}_{2}\mathrm{He}$. We can initially represent the production of an α particle by ${}^{226}_{88}\mathrm{Ra}$ as follows:

$${}^{226}_{88}\mathrm{Ra} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{A}_{Z}\mathrm{X}$$

Because we know that both A and Z are conserved, we can write

$$A + 4 = 226 \quad \text{and} \quad Z + 2 = 88$$

Solving for A gives 222 and for Z gives 86, so ${}^{A}_{Z}\mathrm{X}$ is ${}^{222}_{86}\mathrm{X}$. Because Rn has $Z = 86$, ${}^{A}_{Z}\mathrm{X}$ is ${}^{222}_{86}\mathrm{Rn}$. The overall balanced equation is

$${}^{226}_{88}\mathrm{Ra} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{222}_{86}\mathrm{Rn}$$

Check: $Z = 88 \quad \rightarrow \quad Z = 86 + 2 = 88$
$A = 226 \quad \rightarrow \quad A = 222 + 4 = 226$

b. Using a similar strategy, we have

$${}^{214}_{82}\mathrm{Pb} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{A}_{Z}\mathrm{X}$$

Because $Z - 1 = 82$, $Z = 83$, and because $A + 0 = 214$, $A = 214$. Therefore, ${}^{A}_{Z}\mathrm{X} = {}^{214}_{83}\mathrm{Bi}$. The balanced equation is

$${}^{214}_{82}\mathrm{Pb} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{214}_{83}\mathrm{Bi}$$

Check: $Z = 82 \quad \rightarrow \quad Z = 83 - 1 = 82$
$A = 214 \quad \rightarrow \quad A = 214 + 0 = 214$

Self-Check Exercise 18.2

a. The missing particle must be ${}^{4}_{2}\mathrm{H}$ (an α particle), because

$${}^{222}_{86}\mathrm{Rn} \rightarrow {}^{218}_{84}\mathrm{Po} + {}^{4}_{2}\mathrm{He}$$

is a balanced equation

Check: $Z = 86 \quad \rightarrow \quad Z = 84 + 2 = 86$
$A = 222 \quad \rightarrow \quad A = 218 + 4 = 222$

b. The missing species must be ${}^{15}_{7}\mathrm{X}$ or ${}^{15}_{7}\mathrm{N}$, because the balanced equation is

$${}^{15}_{8}\mathrm{O} \rightarrow {}^{15}_{7}\mathrm{N} + {}^{0}_{1}\mathrm{e}$$

Check: $Z = 8 \quad \rightarrow \quad Z = 7 - 1 = 8$
$A = 15 \quad \rightarrow \quad A = 15 + 0 = 15$

Self-Check Exercise 18.3

Let's do this problem by thinking about the number of half-lives required to go from 8.0×10^{-7} mole to 1.0×10^{-7} mole of ${}^{228}_{88}\mathrm{Ra}$.

$$8.0 \times 10^{-7}\ \text{mol} \xrightarrow[\text{half-life}]{\text{First}} 4.0 \times 10^{-7}\ \text{mol} \xrightarrow[\text{half-life}]{\text{Second}}$$

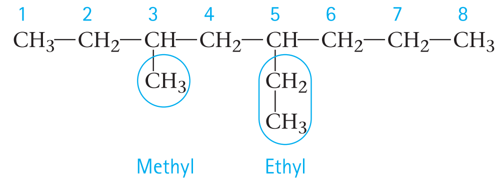
$$2.0 \times 10^{-7}\ \text{mol} \xrightarrow[\text{half-life}]{\text{Third}} 1.0 \times 10^{-7}\ \text{mol}$$

It takes three half-lives, then, for the sample to go from 8.0×10^{-7} mole of ${}^{228}_{88}\mathrm{Ra}$ to 1.0×10^{-7} mole of ${}^{228}_{88}\mathrm{Ra}$. From Table 19.3, we know that the half-life of ${}^{228}_{88}\mathrm{Ra}$ is 6.7 years. Therefore, the elapsed time is 3(6.7 years) = 20.1 years, or 2.0×10^{1} years when we use the correct number of significant figures.

제 19 장

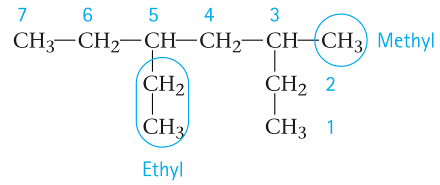
Self-Check Exercise 19.1

a.



This molecule is 5-ethyl-3-methyloctane.

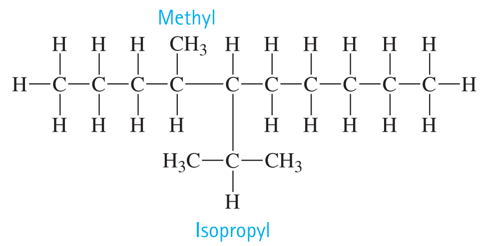
b.



This molecule is 5-ethyl-3-methylheptane. Note that this chain could be numbered from the opposite direction to give the name 3-ethyl-5-methylheptane. These two names are equally

Self=Check Exercise 19.2

The root name *decate* indicates a ten-carbon chain. There is a methyl group at the number 4 position and an isopropyl group at the number 5 position. The structural formula is



Self Check Exercise 19.3

a. The longest chain has eight carbon atoms with a double bond, so the root name is octene. The double bond exists between carbons 3 and 4, so the name is 3-octene. There is a methyl group on the number-2 carbon. The name is 2-methyl-3-octene.

b. The carbon chain has five carbons with a triple bond between carbons 1 and 2. The name is 1-pentyne.

Self-Check Exercise 19.4

a. 2-chloronitrobenzene or *o*-chloronitrobenzene

b. 4-phenyl-2-hexene

Self-Check Exercise 19.5

a. 1-pentanol; primary alcohol

b. 2-methyl-2-propanol (but this alcohol is usually called tertiary butyl alcohol); tertiary alcohol

c. 5-bromo-2-hexanol; secondary alcohol

Self-Check Exercise 19.6

a. 4-ethyl-3-hexanone
Because the compound is named as a hexanone, the carbonyl group is assigned the lowest possible number.

b. 7-isopropyldecanal

▸ 질문과 문제의 선택된 번호 정답

제1장

4. Answers will depend on the student's choices.

5. Recognize the problem and state it clearly; propose possible solutions or explanations; decide which solution/explanation is best through experments.

7. Answers will depend on student responses. A quantitative observation must include a number, such as "There are three windows in this room." A qualitative observation could include something like "The chair is blue."

8. A natural law is a *summary of observed, measurable behavior* that occurs repeatedly and consistently. A theory is an attempt to explain such behavior.

9. Chemistry is not just a set of facts that have to be memorized. To be successful in chemistry, you have to be able to apply what you have learned to new situations, new phenomena, new experiments. Rather than just learning a list of facts or studying someone else's solution to a problem, your instructor hopes you will learn *how* to solve problems *yourself*, so that you will be able to apply what you have learned in future circumstances.

제2장

4. time

5. (a) 10^{-2}; (b) 10^{6}; (c) 10^{9}; (d) 10^{-1}; (e) 10^{-3}; (f) 10^{-6}

7. the woman

9. (a) centimeter; (b) meter; (c) kilometer

10. uncertainty

13. (a) one; (b) infinite (definition); (c) infinite (fixed number); (d) two; (e) two

14. (a) 9.96×10^{-1}; (b) 4.40×10^{3}; (c) 8.22×10^{-1}; (d) 4.00×10^{-9}; (e) 8.42×10^{-2}

15. (a) 8.8×10^{-4}; (b) 9.375×10^{4}; (c) 8.97×10^{-1}; (d) 1.00×10^{3}

16. (a) 5.16; (b) 2423 (2.423×10^{3}); (c) 0.516 (5.16×10^{-1}); (d) 2423

17. (a) one; (b) four; (c) two; (d) three

18. (a) 2.045; (b) 3.8×10^{3}; (c) 5.19×10^{-5}; (d) 3.8418×10^{-7}

20. $\frac{1000\text{ mL}}{1\text{ L}}$; $\frac{1\text{ L}}{1000\text{ mL}}$

21. (a) 2.44 yd; (b) 42.2 m; (c) 115 in.; (d) 2238 cm; (e) 648.1 mi; (f) 716.9 km; (g) 0.0362 km; (h) 5.01×10^{4} cm **23.** 1×10^{-8} cm; 4×10^{-9} in.; 0.1 nm

24. freezing/melting **26.** 273 **27.** Fahrenheit (F)

28. (a) 63 K; (b) 2 °C; (c) 505 °C; (d) 1051 K

29. (a) 173 °F; (b) 104 °F; (c) −459 °F; (d) 90. °F

31. g/cm^3 (g/mL) **33.** copper

34. (a) 22 g/cm^3; (b) 0.034 g/cm^3; (c) 0.962 g/cm^3; (d) 2.1×10^{-5} g/cm^3

35. 0.91 L (two significant figures)

36. 11.7 mL

37. (a) 966 g; (b) 394 g; (c) 567 g; (d) 135 g

제3장

2. forces

4. liquids

6. gaseous

9. Because gases are mostly empty space, they can be *compressed* easily to smaller volumes. In solids and liquids, most of the sample's bulk volume is filled with the molecules, leaving little empty space.

11. chemical

13. malleable; ductile

14. (a) physical; (b) chemical; (c) chemical; (d) chemical; (e) physical; (f) physical; (g) chemical; (h) physical; (i) physical; (j) physical; (k) chemical

16. Compounds consist of two or more elements combined together chemically in a fixed composition, no matter what their source may be. For example, water on earth consists of molecules containing one oxygen atom and two hydrogen atoms. Water on Mars (or any other planet) has the same composition.

19. Assuming the magnesium and sulfur had been measured out in *exactly the correct ratio for complete reaction*, what would remain after heating would be a pure compound. If there were an *excess* of either magnesium or sulfur, however, the material left after reaction would be a *mixture* of the compound and the excess reagent.

21. Heterogeneous mixtures: salad dressing, jelly beans, the change in my pocket; solutions: window cleaner, shampoo, rubbing alcohol

22. (a) primarily a pure compound, but fillers and anticaking agents may have been added; (b) mixture; (c) mixture; (d) pure substance

23. (a) homogeneous; (b) heterogeneous; (c) heterogeneous; (d) heterogeneous; (e) homogeneous

25. Consider a mixture of salt (sodium chloride) and sand. Salt is soluble in water; sand is not. The mixture is added to water and stirred to dissolve the salt, and is then filtered. The salt solution passes through the filter; the sand remains on the filter. The water can then be evaporated from the salt.

제4장

2. Robert Boyle

4. 115 elements are known; 88 occur naturally; 27 are manmade. Table 4.1 lists the most common elements on the earth.

6. The symbols for these elements are based upon their names in other languages.

7. (a) 8; (b) 5; (c) 2; (d) 9; (e) 13; (f) 12; (g) 6; (h) 11; (i) 7; (j) 1

8. praseodymium, Pr; lawrencium, Lr; californium, Cf; nobelium, No; hafnium, Hf

10. (a) Elements are made of tiny particles called atoms. (b) All atoms of a given element are identical; (c) The atoms of a given element are different from those of any other element; (d) A given compound always has the same numbers and types of atoms; (e) Atoms are neither created nor destroyed in chemical processes. A chemical reaction simply changes the way the atoms are grouped together.

12. (a) PbO_2; (b) $CoCl_3$; (c) $C_6H_{12}O_6$; (d) Al_2O_3; (e) Na_2CO_3; (f) CaH_2

13. (a) False; Rutherford's bombardment experiments with metal foil suggested that the α particles were being deflected by coming near a *dense*, *positively charged* atomic nucleus; (b) False; the proton and the electron have opposite charges, but the mass of the electron is *much smaller* than the mass of the proton; (c) True

15. The protons and neutrons are found in the nucleus. The protons are positively charged; the neutrons have no charge. The protons and neutrons each weigh approximately the same.

17. neutron; electron

18. The electrons; outside the nucleus

20. mass

22. Atoms of the same element (atoms with the same number of protons in the nucleus) may have different numbers of neutrons, and so will have different masses.

23.

Z	Symbol	Name
14	Si	silicon
54	Xe	xenon
79	Au	gold
56	Ba	barium
53	I	iodine
50	Sn	tin
48	Cd	cadmium

24. (a) ${}^{26}_{14}Si$; (b) ${}^{30}_{15}P$; (c) ${}^{47}_{24}Cr$; (d) ${}^{60}_{27}Co$; (e) ${}^{62}_{30}Zn$; (f) ${}^{39}_{19}K$

25. (a) 19 protons, 20 neutrons, 19 electrons;
(b) 24 protons, 29 neutrons, 24 electrons;
(c) 34 protons, 50 neutrons, 34 electrons;
(d) 33 protons, 43 neutrons, 33 electrons;
(e) 36 protons, 55 neutrons, 36 protons;
(f) 27 protons, 32 neutrons, 27 electrons

26.

Name	Neutrons	Atomic Number	Mass Number	Symbol
nitrogen	6	7	13	${}^{13}_{7}N$
nitrogen	7	7	14	${}^{14}_{7}N$
lead	124	82	206	${}^{206}_{82}Pb$
iron	31	26	57	${}^{57}_{26}Fe$
krypton	48	36	84	${}^{84}_{36}Kr$

29. Metallic elements are found toward the *left* and bottom of the periodic table; there are far more metallic elements than nonmetals.

31. nonmetallic gaseous elements: oxygen, nitrogen, fluorine, chlorine, hydrogen, and the noble gases; There are no metal- lic gaseous elements at room conditions

33. A metalloid is an element that has some properties common to both metallic and nonmetallic elements. The metalloids are found in the "stair-step" region marked on most periodic tables.

35. (a) fluorine, chlorine, bromine, iodine, astatine;
(b) lithium, sodium, potassium, rubidium, cesium, francium;
(c) beryllium, magnesium, calcium, strontium, barium, radium; (d) helium, neon, argon, krypton, xenon, radon

36.

Name	Neutrons	Atomic Number	Mass Number	Symbol
calcium	Ca	20	2	metal
radon	Rn	86	8	nonmetal
rubidium	Rb	37	1	metal
phosphorus	P	15	5	nonmetal
germanium	Ge	32	4	metalloid

38. These elements are found uncombined in nature and do not readily react with other elements. Although these elements were once thought to form no compounds, this now has been shown to be untrue.

39. diatomic gases: H_2, N_2, O_2, F_2, Cl_2; monatomic gases: He, Ne, Kr, Xe, Rn, Ar

40. chlorine **42.** diamond **44.** electrons

46. 3+ **48.** *-ide* **50.** nonmetallic

51. (a) 10; (b) 22; (c) 10; (d) 10; (e) 23; (f) 54; (g) 23; (h) 2

53. (a) two electrons gained; (b) three electrons gained; (c) three electrons lost; (d) two electrons lost; (e) one electron lost; (f) two electrons lost

54. (a) P^{3-}; (b) Ra^{2+}; (c) At^-; (d) no ion; (e) Cs^+; (f) Se^{2-}

55. Sodium chloride is an *ionic* compound, consisting of Na^+ and Cl^- *ions*. When NaCl is dissolved in water, these ions are *set free* and can move independently to conduct the electric current. Sugar crystals, although they may *appear* similar visually, contain *no* ions. When sugar is dissolved in water, it dissolves as uncharged *molecules*. No electrically charged species are present in a sugar solution to carry the electric current.

57. (a) Cr_2S_3; (b) CrO; (c) AlF_3; (d) Al_2O_3; (e) AlP; (f) Li_3N

제5장

2. A binary chemical compound contains only two elements; the major types are ionic (compounds of a metal and a nonmetal) and nonionic or molecular (compounds between two nonmetals). Answers depend on student responses.

4. Some substances do not contain molecules; the formula we write reflects only the relative number of each type of atom present.

5. Roman numeral

6. (a) potassium bromide; (b) zinc chloride; (c) cesium oxide; (d) magnesium sulfide; (e) aluminum iodide; (f) magnesium bromide; (g) beryllium fluoride; (h) barium hydride

8. (a) Ag_2S; (b) BaH_2; (c) Al_2O_3; (d) MgF_2; (e) correct

9. (a) copper(II) chloride; (b) copper(I) iodide;
(c) manganese(II) bromide; (d) chromium(II) iodide;
(e) chromium(III) chloride; (f) mercury(II) oxide

10. (a) cupric iodide; (b) mercurous bromide;
(c) chromous bromide; (d) cobaltous oxide;
(e) cobaltic oxide; (f) stannous chloride

11. (a) xenon difluoride; (b) diboron trisulfide;
(c) dichlorine hept(a)oxide; (d) silicon tetrabromide;
(e) nitrogen monoxide; (f) sulfur trioxide

12. (a) barium nitride; (b) aluminum sulfide; (c) diphosphorus trisulfide; (d) calcium phosphide; (e) krypton pentafluoride; (f) copper(I) selenide/cuprous selenide

13. (a) barium fluoride; (b) radium oxide; (c) dinitrogen oxide; (d) rubidium oxide; (e) diarsenic pent(a)oxide; (f) calcium nitride

15. An oxyanion is a polyatomic ion containing a given element and one or more oxygen atoms. The oxyanions of chlorine and bromine are given below:

Oxyanion	Name	Oxyanion	Name
ClO^-	hypochlorite	BrO^-	hypobromite
ClO_2^-	chlorite	BrO_2^-	bromite
ClO_3^-	chlorate	BrO_3^-	bromate
ClO_4^-	perchlorate	BrO_4^-	perbromate

17. (a) NO_3^-; (b) NO_2^-; (c) NH_4^+; (d) CN^-

18. (a) ammonium ion; (b) dihydrogen phosphate ion;
(c) sulfate ion; (d) hydrogen sulfite ion (bisulfite ion);
(e) perchlorate ion; (f) iodate ion

19. (a) ammonium acetate; (b) lithium perchlorate; (c) sodium hydrogen sulfate; (d) gold(III) carbonate; (e) calcium chlorate; (f) hydrogen peroxide

21. (a) hypochlorous acid; (b) sulfurous acid; (c) bromic acid; (d) hypoiodous acid; (e) perbromic acid; (f) hydrosulfuric acid; (g) hydroselenic acid; (h) phosphorous acid

42. (a) lithium nitrate; (b) chromium(III) carbonate/chromic carbonate; (c) copper(II) carbonate/cupric carbonate; (d) copper (I) selenide/cuprous selenide; (e) manganese(IV) sulfate; (f) magnesium nitrite

24. (a) N_2O; (b) NO_2; (c) N_2O_4; (d) SF_6; (e) PBr_3; (f) CI_4; (g) OCl_2

25. (a) $BaSO_3$; (b) $Ca(H_2PO_4)_2$; (c) NH_4ClO_4; (d) $NaMnO_4$; (e) $Fe_2(SO_4)_3$; (f) $CoCO_3$; (g) $Ni(OH)_2$; (h) $ZnCrO_4$

26. (a) HCN; (b) HNO_3; (c) H_2SO_4; (d) H_3PO_4; (e) HClO or HOCl; (f) HBr; (g) $HBrO_2$; (h) HF

27. (a) $Mg(HSO_4)_2$; (b) $CsClO_4$; (c) FeO; (d) $H_2Te(aq)$; (e) $Sr(NO_3)_2$; (f) $Sn(C_2H_3O_2)_4$; (g) $MnSO_4$; (h) N_2O_4; (i) Na_2HPO_4; (j) Li_2O_2; (k) HNO_2; (l) $Co(NO_3)_3$

제6장

2. Most of these products contain a peroxide, which decomposes and releases oxygen gas.

3. Bubbling takes place as the hydrogen peroxide chemically decomposes into water and oxygen gas.

7. water

8. $H_2O_2(aq) \rightarrow H_2(g) + O_2(g)$

9. $AgNO_3(aq) + HCl(aq) \rightarrow AgCl(s) + HNO_3(aq)$; $Pb(NO_3)_2(aq) + HCl(aq) \rightarrow PbCl_2(s) + HNO_3(aq)$

10. $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$; $C_3H_8(g) + O_2(g) \rightarrow CO(g) + H_2O(g)$

11. $CaCO_3(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

12. $SiO_2(s) + C(s) \rightarrow Si(s) + CO(g)$

13. $H_2S(g) + O_2(g) \rightarrow SO_2(g) + H_2O(g)$

14. $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$; $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

15. $NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$

16. $NH_3(g) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$

17. $Xe(g) + F_2(g) \rightarrow XeF_4(s)$

18. $NH_4Cl(s) + NaOH(s) \xrightarrow{heat} NH_3(g) + H_2O(g) + NaCl(s)$

20. whole numbers

21. (a) $2Al(s) + 3CuO(s) \rightarrow Al_2O_3(s) + 3Cu(l)$; (b) $S_8(s) + 24F_2(g) \rightarrow 8SF_6(g)$; (c) $Xe(g) + 3F_2(g) \rightarrow XeF_6(s)$; (d) $NH_4Cl(g) + KOH(s) \rightarrow NH_3(g) + H_2O(g) + KCl(s)$; (e) $SiC(s) + 2Cl_2(g) \rightarrow SiCl_4(l) + C(s)$; (f) $K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$; (g) $N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(aq)$; (h) $8H_2S(g) + 8Cl_2(g) \rightarrow S_8(s) + 16HCl(g)$

22. (a) $Na_2SO_4(aq) + CaCl_2(aq) \rightarrow CaSO_4(s) + 2NaCl(aq)$; (b) $3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$; (c) $Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$; (d) $Br_2(g) + 2H_2O(l) + SO_2(g) \rightarrow 2HBr(aq) + H_2SO_4(aq)$; (e) $3NaOH(s) + H_3PO_4(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$; (f) $2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$; (g) $2Na_2O_2(s) + 2H_2O(l) \rightarrow 4NaOH(aq) + O_2(g)$; (h) $4Si(s) + S_8(s) \rightarrow 2Si_2S_4(s)$

23. (a) $4NaCl(s) + 2SO_2(g) + 2H_2O(g) + O_2(g) \rightarrow 2Na_2SO_4(s) + 4HCl(g)$; (b) $3Br_2(l) + I_2(s) \rightarrow 2IBr_3(s)$; (c) $Ca(s) + 2H_2O(g) \rightarrow Ca(OH)_2(aq) + H_2(g)$; (d) $2BF_3(g) + 3H_2O(g) \rightarrow B_2O_3(s) + 6HF(g)$; (e) $SO_2(g) + 2Cl_2(g) \rightarrow SOCl_2(l) + Cl_2O(g)$; (f) $Li_2O(s) + H_2O(l) \rightarrow 2LiOH(aq)$; (g) $Mg(s) + CuO(s) \rightarrow MgO(s) + Cu(l)$; (h) $Fe_3O_4(s) + 4H_2(g) \rightarrow 3Fe(l) + 4H_2O(g)$

24. (a) $Ba(NO_3)_2(aq) + Na_2CrO_4(aq) \rightarrow BaCrO_4(s) + 2NaNO_3(aq)$; (b) $PbCl_2(aq) + K_2SO_4(aq) \rightarrow PbSO_4(s) + 2KCl(aq)$; (c) $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$; (d) $CaC_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$; (e) $Sr(s) + 2HNO_3(aq) \rightarrow Sr(NO_3)_2(aq) + H_2(g)$; (f) $BaO_2(s) + H_2SO_4(aq) \rightarrow BaSO_4(s) + H_2O_2(aq)$; (g) $2AsI_3(s) \rightarrow 2As(s) + 3I_2(s)$; (h) $2CuSO_4(aq) + 4KI(s) \rightarrow 2CuI(s) + I_2(s) + 2K_2SO_4(aq)$

제7장

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

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

7. The simplest evidence is that solutions of ionic substances conduct electricity.

9. b, c, f, h

10. (a) Rule 5; (b) Rule 6; (c) Rule 6; (d) Rule 6

11. (a) $MnCO_3$, Rule 6; (b) $CaSO_4$, Rule 4; (c) Hg_2Cl_2, Rule 3; (d) no precipitate, most sodium and nitrate salts are soluble; (e) $Ni(OH)_2$, Rule 5; (f) $BaSO_4$, Rule 4

12. (a) $Na_2S(aq) + CuCl_2(aq) \rightarrow \underline{CuS}(s) + 2NaCl(aq)$; (b) $K_3PO_4(aq) + AlCl_3(aq) \rightarrow \underline{AlPO_4}(s) + 3KCl(aq)$; (c) $H_2SO_4(aq) + BaCl_2(aq) \rightarrow \underline{BaSO_4}(s) + 2HCl(aq)$; (d) $3NaOH(aq) + FeCl_3(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$; (e) $2NaCl(aq) + Hg_2(NO_3)_2(aq) \rightarrow \underline{Hg_2Cl_2}(s) + 2NaNO_3(aq)$; (f) $3K_2CO_3(aq) + 2Cr(C_2H_3O_2)_3(aq) \rightarrow \underline{Cr_2(CO_3)_3}(s) + 6KC_2H_3O_2(aq)$

13. (a) $CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$; (b) $2AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2KNO_3(aq)$; (c) $BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KCl(aq)$

14. (a) $CaCl_2(aq) + 2AgC_2H_3O_2(aq) \rightarrow \mathbf{2AgCl}(s) + Ca(C_2H_3O_2)_2(aq)$; (b) $Ba(NO_3)_2(aq) + 2NH_4OH(aq) \rightarrow \mathbf{Ba(OH)_2}(s) + 2NH_4NO_3(aq)$; (c) $NiCl_2(aq) + Na_2CO_3(aq) \rightarrow \mathbf{NiCO_3}(s) + 2NaCl(aq)$

16. 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 spectator ions.

17. (a) $Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$; (b) $Ni^{2+}(aq) + 2OH^-(aq) \rightarrow Ni(OH)_2(s)$; (c) $2Fe^{3+}(aq) + 3S^2(aq) \rightarrow Fe_2S_3(s)$

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

20. The strong bases are those hydroxide compounds that dissociate fully when dissolved in water. The strong bases that are highly soluble in water (NaOH, KOH) are also strong electrolytes.

23. 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.

25. $RbOH(s) \rightarrow Rb^+(aq) + OH^-(aq)$; $CsOH(s) \rightarrow Cs^+(aq) + OH^-(aq)$

27. (a) $HCl(aq) + KOH(aq) \rightarrow H_2O(l) + KCl(aq)$; (b) $HClO_4(aq) + NaOH(aq) \rightarrow NaClO_4(aq) + H_2O(l)$; (c) $CsOH(aq) + HNO_3(aq) \rightarrow CsNO_3(aq) + H_2O(l)$; (d) $2KOH(aq) + H_2SO_4(aq) \rightarrow 2H_2O(l) + K_2SO_4(aq)$

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

30. The metal loses electrons, the nonmetal gains electrons.

31. 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 O_2 molecule. Magnesium ions are charged 2+, oxide ions are charged 2−.

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

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

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

33. (a) $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$; (b) $MgO(s) + C(s) \rightarrow Mg(s) + CO(g)$; (c) $Sr(s) + 2H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$; (d) $Co(s) + 2HCl(aq) \rightarrow CoCl_2(aq) + H_2(g)$
35. (a) oxidation–reduction; (b) oxidation–reduction; (c) acid–base; (d) acid–base, precipitation; (e) precipitation; (f) precipitation; (g) oxidation–reduction; (h) oxidation–reduction; (i) acid–base
37. oxidation–reduction
39. A decomposition reaction is one in which a given compound is broken down into simpler compounds or constituent elements. The reactions $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ and $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$ represent decomposition reactions. Such reactions often may be classified in other ways. For example, the reaction of $HgO(s)$ is also an oxidation–reduction reaction.
41. (a) $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$; (b) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$; (c) $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$
42. (a) $8Fe(s) + S_8(s) \rightarrow 8FeS(s)$; (b) $4Co(s) + 3O_2(g) \rightarrow 2Co_2O_3(s)$; (c) $Cl_2O_7(g) + H_2O(l) \rightarrow 2HClO_4(aq)$
43. (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) $3NaN_3(s) \rightarrow Na_3N(s) + 4N_2(g)$

제8장

2. 307 corks; 116 stoppers; 613 corks; 2640 g (2.64×10^3 g)
4. The average atomic mass takes into account the various isotopes of an element and the relative abundances in which those isotopes are found.
5. (a) one; (b) five; (c) ten; (d) 50; (e) ten
6. A sample containing 35 tin atoms would weigh 4155 amu; 2967.5 amu of tin would represent 25 tin atoms.
7. $3 \times$ Avogadro's number ($3 \times 6.022 \times 10^{23} = 1.807 \times 10^{24}$)
8. 32.00 g
9. 177 g
10. 2.326×10^{-23} g
11. 0.50 mol O
12. (a) 3.500 mol of F atoms; (b) 2.000 mmol Hg; (c) 3.000 mol Si; (d) 0.2500 mol Pt; (e) 100.0 mol Mg; (f) 0.5000 mol Mo
13. (a) 2.34×10^{-1} g; (b) 0.252 g; (c) 1.10×10^6 g; (d) 2.80×10^{-5}g; (e) 0.413 g; (f) 105 g
14. (a) 9.77×10^3 amu; (b) 1.62×10^{-20} g; (c) 9.77×10^3 g; (d) 2.56×10^{26} atoms; (e) 1.11×10^{25} atoms; (f) 18.5 mol Na; (g) 449 g Mg
16. The molar mass is calculated by summing the individual atomic masses of the atoms in the formula.
17. (a) carbon monoxide, 28.01 g; (b) sodium carbonate, 105.99 g; (c) iron(III) nitrate/ferric nitrate, 241.88 g; (d) hydrogen iodide, 127.9 g; (e) sulfur trioxide, 80.07 g
18. (a) aluminum fluoride, 83.98 g; (b) sodium phosphate, 163.94 g; (c) magnesium carbonate, 84.32 g; (d) lithium hydrogen carbonate/lithium bicarbonate, 67.96 g; (e) chromium(III) oxide/chromic oxide, 152 g
19. (a) 8.92×10^{-4} mol NaCl; (b) 0.125 mol $MgCO_3$; (c) 0.0392 mol Al_2O_3; (d) 0.151 mol Fe_2O_3; (e) 1.69×10^{-3} mol Li_2CO_3; (f) 40.3 mol Fe
20. (a) 3.55×10^{-5} mol; (b) 5.26 mol; (c) 2704 mol; (d) 0.0697 mol; (e) 467 mol
21. (a) 0.132 g; (b) 5.31×10^{-4} g; (c) 3.70×10^4 g; (d) 0.633 g; (e) 0.115 g; (f) 112 g
22. (a) 0.313 g; (b) 139 g; (c) 1.67 g; (d) 11.3 g; (e) 4.66×10^4 g; (f) 4.45 g
23. (a) 3.84×10^{24} molecules; (b) 1.37×10^{23} molecules; (c) 8.76×10^{16} molecules; (d) 1.58×10^{18} molecules; (e) 4.03×10^{22} molecules
24. (a) 0.0141 mol S; (b) 0.0159 mol S; (c) 0.0258 mol S; (d) 0.0127 mol S
26. less
27. (a) 88.82% Cu, 11.18% O; (b) 79.89% Cu, 20.11% O; (c) 77.73% Fe, 22.27% O; (d) 69.94% Fe, 30.06% O; (e) 46.68% N, 53.32% O; (f) 30.45% N, 69.55% O
28. (a) 28.45% Cu; (b) 44.29% Cu; (c) 44.06% Fe; (d) 34.43% Fe; (e) 18.85% Co; (f) 13.40% Co; (g) 88.12% Sn; (h) 78.77% Sn
29. (a) 34.43% Fe; (b) 29.63% O; (c) 92.25% C; (d) 11.92% N; (e) 93.10% Ag; (f) 45.39% Co; (g) 30.45% N; (h) 43.66% Mn
33. a, c
34. BaH_2
35. $CaO_2Cl_2/Ca(OCl)_2$
36. $N_2H_8CO_3$ [actually $(NH_4)_2CO_3$]
37. Co_2S_3
38. AlF_3
39. AlF_3
40. Li_3N
41. $Al_2S_3O_{12}$ [actually $Al_2(SO_4)_3$]
42. PCl_3, PCl_5
44. molar mass
46. $C_4H_{10}O_2$

제9장

2. The coefficients of the balanced chemical equation indicate the relative numbers of molecules (or moles) of each reactant that combine, as well as the number of molecules (or moles) of each product formed.
3. Balanced chemical equations tell us in what molar ratios substances combine to form products, not in what mass proportions they combine.
4. (a) $(NH_4)_2CO_3(s) \rightarrow 2NH_3(g) + CO_2(g) + H_2O(g)$.
   One formula unit of solid ammonium carbonate decomposes to produce two molecules of ammonia gas, one molecule of carbon dioxide gas, and one molecule of water vapor. One mole of solid ammonium carbonate decomposes into two moles of gaseous ammonia, one mole of carbon dioxide gas, and one mole of water vapor.
   (b) $6Mg(s) + P_4(s) \rightarrow 2Mg_3P_2(s)$.
   Six atoms of magnesium metal react with one molecule of solid phosphorus (P_4) to make two formula units of solid magnesium phosphide. Six moles of magnesium metal react with one mole of solid phosphorus (P_4) to produce two moles of solid magnesium phosphide.
   (c) $4Si(s) + S_8(s) \rightarrow 2Si_2S_4(l)$.
   Four atoms of solid silicon react with one molecule of solid sulfur (S_8) to form two molecules of liquid disilicon tetrasulfide. Four moles of solid silicon react with one mole of solid sulfur (S_8) to form two moles of liquid disilicon tetrasulfide.
   (d) $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$.
   One molecule of liquid ethanol burns with three molecules of oxygen gas to produce two molecules of carbon dioxide gas and three molecules of water vapor. One mole of liquid ethanol burns with three moles of oxygen gas to produce two moles of gaseous carbon dioxide and three moles of water vapor.

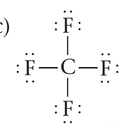
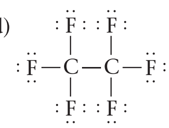
5. Balanced chemical equations tell us in what molar ratios substances combine to form products, not in what mass proportions they combine. How could 2 g of reactant produce a total of 3 g of products?
6. $S(s) + 2H_2SO_4(l) \rightarrow 3SO_2(g) + 2H_2O(l)$;
for SO_2, $\left(\frac{3 \text{ mol } SO_2}{1 \text{ mol } S}\right)$; for H_2O, $\left(\frac{2 \text{ mol } H_2O}{1 \text{ mol } S}\right)$
for H_2SO_4, $\left(\frac{2 \text{ mol } H_2SO_4}{1 \text{ mol } S}\right)$
7. (a) 17.5 mol/18 mol; (b) 25 mol; (c) 5.0 mol; (d) 3.75/3.8 mol
8. (a) 0.50 mol NH_4Cl (27 g); (b) 0.13 mol CS_2 (9.5 g); 0.25 mol H_2S (8.5 g); (c) 0.50 mol H_3PO_3 (41 g); 1.5 mol HCl (55 g); (d) 0.50 mol $NaHCO_3$ (42 g)
9. (a) 0.469 mol O_2; (b) 0.938 mol Se; (c) 0.625 mol CH_3CHO; (d) 1.25 mol Fe
11. (a) 3.19×10^{-5} mol; (b) 8.07×10^{-6} mol; (c) 11.5 mol; (d) 0.0256 mol; (e) 0.138 mol
12. (a) 119 g; (b) 678 g; (c) 0.0438 g; (d) 256 g; (e) 0.0206 g; (f) 170. g; (g) 2.11×10^{-4} g
13. (a) 0.326 mol; (b) 0.202 mol; (c) 0.124 mol; (d) 0.0448 mol
14. (a) 1.38 g B, 14.0 g HCl; (b) 13.5 g Cu_2O, 6.04 g SO_2; (c) 35.9 g Cu, 6.04 g SO_2; (d) 29.0 g $CaSiO_3$, 11.0 g CO_2
15. 3.82 g H_2SO_4
16. 0.959 g Na_2CO_3
17. 2.68 g ethyl alcohol
18. 0.443 g NH_3
19. 8.62 kg Hg
20. 0.501 g C
21. 2.07 g MgO
22. To determine the limiting reactant, first calculate the number of moles of each reactant present. Then determine how these numbers of moles correspond to the stoichiometric ratio indicated by the balanced chemical equation for the reaction. For each reactant, use the stoichiometric ratios from the balanced chemical equation to calculate how much of the *other* reactants would be required to react completely.
24. (a) H_2SO_4 is limiting, 4.90 g SO_2, 0.918 g H_2O; (b) H_2SO_4 is limiting, 6.30 g $Mn(SO_4)_2$, 0.918 g H_2O; (c) O_2 is limiting, 6.67 g SO_2, 1.88 g H_2O; (d) $AgNO_3$ is limiting, 3.18 g Ag, 2.09 g $Al(NO_3)_3$
25. (a) O_2 is limiting, 0.458 g CO_2; (b) CO_2 is limiting, 0.409 g H_2O; (c) MnO_2 is limiting, 0.207 g H_2O; (d) I_2 is limiting, 1.28 g ICl
26. (a) CO is limiting reactant; 11.4 mg CH_3OH; (b) I_2 is limiting reactant; 10.7 mg AlI_3; (c) HBr is limiting reactant; 12.4 mg $CaBr_2$; 2.23 mg H_2O; (d) H_3PO_4 is limiting reactant; 15.0 mg $CrPO_4$; 0.309 mg H_2
27. 136 g urea
28. 1.79 g Fe_2O_3
29. 0.627 g CuI; 0.691 g KI_3; 0.573 g K_2SO_4
30. 0.67 kg SiC
32. theoretical yield, 1.16 g; percent yield, 94.0%
33. $2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(g)$. 142 g of CO_2 can be ultimately absorbed; 102 g is 71.8% of the canister's capacity.
34. theoretical, 2.72 g $BaSO_4$; percent, 74.3%

제10장

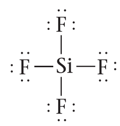
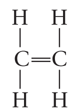
2. potential
4. The total energy of the universe is constant. Energy cannot be created or destroyed, but can only be converted from one form to another.
6. Ball A initially possesses potential energy by virtue of its position at the top of the hill. As ball A rolls down the hill, its potential energy is converted to kinetic energy and frictional (heat) energy. When ball A reaches the bottom of the hill and hits ball B, it transfers its kinetic energy to ball B. Ball A then has only the potential energy corresponding to its new position.
8. The hot tea is at a higher temperature, which means the particles in the hot tea have higher average kinetic energies. When the tea spills on the skin, energy flows from the hot tea to the skin, until the tea and skin are at the same temperature. This sudden inflow of energy causes the burn.
10. Temperature is the concept by which we express the thermal energy contained in a sample. We cannot measure the motions of the particles/kinetic energy in a sample of matter directly. We know, however, that if two objects are at different temperatures, the one with the higher temperature has molecules that have higher average kinetic energies than the molecules of the object at the lower temperature.
12. When the chemical system evolves energy, the energy evolved from the reacting chemicals is transferred to the surroundings.
14. exactly equal to 16. internal
17. losing 18. gaining
20. (a) $\frac{1 \text{ J}}{4.184 \text{ cal}}$; (b) $\frac{4.184 \text{ cal}}{1 \text{ J}}$; (c) $\frac{1 \text{ kcal}}{1000 \text{ cal}}$; (d) $\frac{1000 \text{ J}}{1 \text{ kJ}}$
21. 6540 J = 6.54 kJ
23. (a) 243 kJ to three significant figures; (b) 0.004184 kJ; (c) 0.000251 kJ; (d) 0.4503 kJ
26. 89 cal 27. 29°C 28. 0.057 cal/g °C
29. The specific heat is 0.89 J/g °C, so the element is most likely aluminum.
31. (a) −9.23 kJ; (b) −148 kJ; (c) +296 kJ/mol
32. (a) −445 kJ/mol water; (b) −445 kJ/mol dioxygen
33. −220 kJ
34. −233 kJ
35. Once everything in the universe is at the same temperature, no further thermodynamic work can be done. Even though the total energy of the universe will be the same, the energy will have been dispersed evenly, making it effectively useless.
38. Tetraethyl lead was used as an additive for gasoline to promote smoother running of engines. It is no longer widely used because of concerns about the lead being released into the environment as the leaded gasoline burns.
39. The greenhouse effect is a warming effect due to the presence of gases in the atmosphere that absorb infrared radiation that has reached the earth from the sun; the gases do not allow the energy to pass back into space. A limited greenhouse effect is desirable because it moderates the temperature changes in the atmosphere that would otherwise be more drastic between daytime when the sun is shining and nighttime. Having too high a concentration of greenhouse gases, however, will elevate the temperature of the earth too much, affecting climate, crops, the polar ice caps, the temperatures of the oceans, and so on. Carbon dioxide produced by combustion reactions is our greatest concern as a greenhouse gas.
41. If a proposed reaction involves either or both of those phenomena, the reaction will tend to be favorable.
44. The molecules in liquid water are moving around freely and are therefore more "disordered" than when the molecules are held rigidly in a solid lattice in ice. The entropy increases during melting.

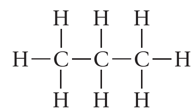
제 11 장

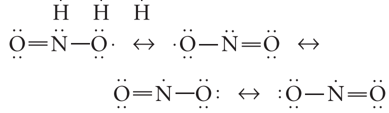
2. bond energy
4. covalent
5. In H_2 and HF, the bonding is covalent in nature, with an electron pair being shared between the atoms. In H_2, the two atoms are identical and so the sharing is equal; in HF, the two atoms are different and so the bonding is polar covalent. Both of these are in marked contrast to the situation in NaF: NaF is an ionic compound, and an electron is completely transferred from sodium to fluorine, thereby producing the separate ions.
7. The difference in electronegativity between the atoms in the bond
8. (a) I is most electronegative, Rb is least electronegative; (b) Mg is most electronegative, Ca and Sr have similar electronegativities; (c) Br is most electronegative, K is least electronegative
9. (a) ionic; (b) polar covalent (c) covalent
10. c and d
11. (a) F; (b) neither; (c) Cl; (d) Cl
12. (a) Ca–Cl; (b) Ba–Cl; (c) Fe–I; (d) Be–F
14. The presence of strong bond dipoles and a large overall dipole moment makes water a very polar substance. Properties of water that are dependent on its dipole moment involve its freezing point, melting point, vapor pressure, and ability to dissolve many substances.
15. (a) H; (b) Cl; (c) I
16. (a) $^{\delta+}P \rightarrow F^{\delta-}$; (b) $^{\delta+}P \rightarrow O^{\delta-}$; (c) $^{\delta+}P \rightarrow C^{\delta-}$; (d) P H (H and P have essentially the same electronegativities)
17. (a) $^{\delta+}H \rightarrow C^{\delta-}$; (b) $^{\delta+}N \rightarrow O^{\delta-}$; (c) $^{\delta+}S \rightarrow N^{\delta-}$; (d) $^{\delta+}C \rightarrow N^{\delta-}$;
19. preceding
21. Atoms in covalent molecules gain a configuration like that of a noble gas by sharing one or more pairs of electrons between atoms: such shared pairs of electrons “belong” to each of the bonding atoms at the same time. In ionic bonding, one atom completely donates one or more electrons to another atom, and then the resulting ions behave independently of one another (they are not “attached” to one another, although they are mutually attracted).
22. (a) Br^- [Kr]; (b) Cs^+ [Xe]; (c) P^{3-} [Ar]; (d) S^{2-} [Ar]
24. (a) Na_2S; (b) BaSe; (c) $MgBr_2$; (d) Li_3N; (e) KH
25. (a) S^{2+} [Kr]; O^{2-} [Ne]; (b) Ca^{2+} [Ar]; H^- [He]; (c) K^+ [Ar]; P^{3-} [Ar]; (d) Ba^{2+} [Xe]; Se^{2-} [Kr]
27. An ionic solid such as NaCl consists of an array of alternating positively and negatively charged ions: that is, each positive ion has as its nearest neighbors a group of negativeions, and each negative ion has a group of positive ions surrounding it. In most ionic solids, the ions are packed as tightly as possible.
29. In forming an anion, an atom gains additional electrons in its outermost (valence) shell. Having additional electrons in the valence shell increases the repulsive forces between elec-trons, and the outermost shell becomes larger to accommo-date this.
30. (a) F^- is larger than Li^+. The F^- ion has a filled n = 2 shell. A lithium atom has lost the electron from its n = 2 shell, leaving the n = 1 shell as its outermost. (b) Cl^- is larger than Na^+, since its valence electrons are in the n = 3 shell (Na^+ has lost its 3s electron). (c) Ca is larger than Ca^{2+}. Positive ions are always smaller than the atoms from which they are formed. (d) I^- is larger. Both Cs^+ and I^- have the same electron configuration (isoelectronic with Xe) and have their valence electrons in the same shell. However, Cs^+ has two more positive charges in its nucleus than does I^-; this charge causes the n = 5 shell of Cs^+ to be smaller than that of I^- (the electrons are pulled in closer to the nucleus by the positive charge).
31. (a) I; (b) F^-; (c) F^-
33. When atoms form covalent bonds, they try to attain a valence-electron configuration similar to that of the following noble gas element. When the elements in the first few horizontal rows of the periodic table form covalent bonds, they attempt to achieve the configurations of the noble gases helium (two valence electrons, duet rule) and neon and argon (eight valence electrons, octet rule).
36. (a) Rb· (b) :Cl· (c) :Kr: (d) Ba: (e) ·P· (f) :At·
37. (a) 8; (b) 16; (c) 26; (d) 26
38. (a) H—H (b) H—Cl:

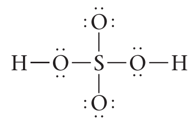
(c)  (d) 

39. (a) H_2S H—S—H

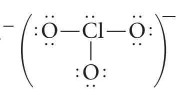
(b) SiF_4  (c) C_2H_4 

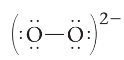
(d) C_3H_8 

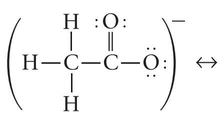
40. (a) NO_2 

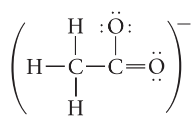
(b) H_2SO_4 

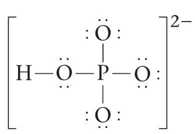
(c) N_2O_4 

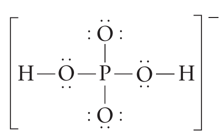
41. (a) ClO_3^- 

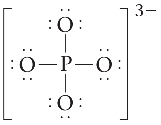
(b) O_2^{2-} 

(c) $C_2H_3O_2^-$ 



42. (a) HPO_4^{2-} 

(b) $H_2PO_4^-$ 

(c) PO_4^{3-} 

44. The geometric structure of NH_3 is that of a trigonal pyramid. The nitrogen atom of NH_3 is surrounded by four electron pairs (three are bonding, one is a lone pair). The H—N—H bond angle is somewhat less than 109.5° (because of the presence of the lone pair).

47. The general molecular structure of a molecule is determined by how many electron pairs surround the central atom in the molecule, and by which of those pairs are used for bonding to the other atoms of the molecule.

49. In NF_3, the nitrogen atom has *four* pairs of valence electrons; in BF_3, only *three* pairs of valence electrons surround the boron atom. The nonbonding pair on nitrogen in NF_3 pushes the three F atoms out of the plane of the N atom.

50. (a) four electron pairs arranged in a tetrahedral arrangement with some distortion due to the nonbonding pair; (b) four electron pairs in a tetrahedral arrangement; (c) four electron pairs in a tetrahedral arrangement

51. (a) tetrahedral; (b) bent(nonlinear); (c) tetrahedral

52. (a) basically tetrahedral arrangement of the oxygens around the phosphorus; (b) tetrahedral; (c) trigonal pyramid

53. (a) approximately tetrahedral (a little less than 109.5°); (b) approximately tetrahedral (a little less than 109.5°); (c) tetrahedral (109.5°); (d) trigonal planar (120°) because of the double bond

54. The ethylene molecule contains a double bond between the carbon atoms. This makes the molecule planar (flat), with H—C—H and H—C—C bond angles of approximately 120°. The 1,2-dibromoethane molecule would *not* be planar, however. Each carbon would h0ave four bonding pairs of electrons around it, and so the orientation around each carbon atom would be basically tetrahedral with bond angles of approximately 109.5° (assuming all bonds are similar).

제12장

2. Solids are rigid and incompressible and have definite shapes and volumes. Liquids are less rigid than solids; although they have definite volumes, liquids take the shape of their containers Gases have no fixed volume or shape; they take the volume and shape of their container and are affected more by changes in their pressure and temperature than are solids or liquids.

3. Pressure units include mm Hg, torr, pascals, and psi. The unit "mm Hg" is derived from the barometer, because in a traditional mercury barometer, we measure the height of the mercury column (in millimeters) above the reservoir of mercury.

4. (a) 0.980 atm; (b) 1.01 atm; (c) 0.916 atm; (d) 3.41×10^{-6} atm

5. (a) 792 mm Hg; (b) 714 mm Hg; (c) 746 mm Hg; (d) 8.18 mm Hg

6. (a) 2.07×10^3 kPa; (b) 106 kPa; (c) 1.10×10^3 kPa; (d) 87.9 kPa

7. Additional mercury increases the pressure on the gas sample, causing the volume of the gas upon which the pressure is exerted to decrease (Boyle's law).

9. $PV = k$; $P_1V_1 = P_2V_2$

10. (a) 4.69 atm; (b) 1.90×10^4 mm Hg; (c) 0.270 atm

11. (a) 146 mL; (b) 0.354 L; (c) 687 mm Hg

12. 0.520 L

13. 27.2 atm

14. Charles's law indicates that an ideal gas decreases by 1/273 of its volume for every Celsius degree its temperature is lowered. This means an ideal gas would approach a volume of zero at −273 °C.

15. $V = bT$; $V_1/T_1 = V_2/T_2$

16. 35.4 mL

17. (a) 80.2 mL; (b) −77 °C (196 K); (c) 208 mL (2.1×10^2 mL)

18. (a) 35.4 K = −238 °C; (b) 0 mL (absolute zero; a real gas would condense to a solid or liquid); (c) 40.5 mL

19. If the temperature is decreased by a factor of 2, the volume will also decrease by a factor of 2. The new volume of the sample will be half its original volume.

20. 90 °C, 124 mL; 80 °C, 121 mL; 70 °C, 117 mL; 60 °C, 113 mL; 50 °C, 110. mL; 40 °C, 107 mL; 30 °C, 103 mL; 20 °C, 99.8 mL

21. $V = an$; $V_1/n_1 = V_2/n_2$

22. 5.60 L

23. 80.1 L

25. Real gases behave most ideally at relatively high temperatures and relatively low pressures. We usually assume that a real gas's behavior approaches ideal behavior if the temper- ature is over 0 °C (273 K) and the pressure is 1 atm or lower.

27. For an ideal gas, $PV = nRT$ is true under any conditions. Consider a particular sample of gas (n remains constant) at a particular fixed pressure (P remains constant). Suppose that at temperature T_1 the volume of the gas sample is V_1. For this set of conditions, the ideal gas equation would be given by $PV_1 = nRT_1$. If the temperature of the gas sample changes to a new temperature, T_2, then the volume of the gas sample changes to a new volume, V_2. For this new set of conditions, the ideal gas equation would be given by $PV_2 = nRT_2$. If we make a ratio of these two expressions for the ideal gas equation for this gas sample, and cancel out terms that are constant for this situation (P, n, and R), we get

$$\frac{PV_1}{\mathrm{PV}_2} = \frac{nRT_1}{nRT_2}, \text{ or } \frac{V_1}{V_2} = \frac{T_1}{T_2},$$

which can be rearranged to the familiar form of Charles's law,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}.$$

28. (a) 5.02 L; (b) 3.56 atm = 2.70×10^3 mm Hg; (c) 334 K

29. 2428 K/2.43×10^3 K

30. 131 atm

31. 0.150 atm; 0.163 atm

32. 238 K/−35 °C **33.** 340 atm

34. 0.332 atm; 0.346 atm **35.** 150. atm

36. As a gas is bubbled through water, the bubbles of gas become saturated with water vapor, thus forming a gaseous mixture. The total pressure for a sample of gas that has been collected by bubbling through water is made up of two components: the pressure of the sample gas and the pressure of water vapor. The partial pressure of the gas equals the total pressure of the sample minus the vapor pressure of water.

37. 0.314 atm

38. 0.0984 mol; 0.0984 mol

39. 3.07 atm

40. $P_{\text{hydrogen}} = 0.990$ atm; 9.55×10^{-3} mol H_2; 0.625 g Zn

42. pressure

43. If the temperature of a sample of gas is increased, the average kinetic energy of the particles of gas increases. This means that the speeds of the particles increase. If the particles have a higher speed, they hit the walls of the container more frequently and with greater force, thereby increasing the pressure.

44. STP = 0 °C, 1 atm pressure. These conditions were chosen because they are easy to attain and reproduce *experimentally*.

The barometric pressure within a laboratory will usually be near 1 atm, and 0 °C can be attained with a simple ice bath.

45. 1.93 L

46. 0.941 L

47. 0.941 L; 0.870 L

48. 5.03 L (dry volume)

49. 52.7 L

50. 184 mL

51. 40.5 L; $P_{He} = 0.864$ atm; $P_{Ne} = 0.136$ atm

52. 1.72 L

53. 0.365 g

제 13 장

3. Because it requires so much more energy to vaporize water than to melt ice, this suggests that the gaseous state is significantly different from the liquid state, but that the liquid and solid states are relatively similar.

4. See Figure 13.2.

6. When a solid is heated, the molecules begin to vibrate/move more quickly. When enough energy has been added to overcome the intermolecular forces that hold the molecules in a crystal lattice, the solid melts. As the liquid is heated, the molecules begin to move more quickly and more randomly. When enough energy has been added, molecules having sufficient kinetic energy will begin to escape from the surface of the liquid. Once the pressure of vapor coming from the liquid is equal to the pressure above the liquid, the liquid boils. Only intermolecular forces need to be overcome in this process: no chemical bonds are broken.

8. intramolecular; intermolecular

9. fusion; vaporization

10. (a) more energy is required to separate the molecules of a liquid into the freely moving and widely separated molecules of a vapor/gas; (b) 0.113 kJ; (c) 4.22 kJ; (d) −4.22 kJ

11. 8.35 kJ; 84.4 kJ; 23.1 kJ

12. 107 kJ/mol; 39.5 kJ; 1.07×10^3 kJ

16. The hydrogen bonding that can exist when H is bonded to O (or N or F) is an additional intermolecular force, which means additional energy must be added to separate the molecules during boiling.

17. London dispersion forces are instantaneous dipole forces that arise when the electron cloud of an atom is momentarily distorted by a nearby dipole, temporarily separating the centers of positive and negative charge in the atom.

18. (a) London dispersion forces; (b) hydrogen bonding; (c) London dispersion forces (d) dipole–dipole forces

20. For a homogeneous mixture to form, the forces between molecules of the two substances being mixed must be at least *comparable in magnitude* to the intermolecular forces within each separate substance. In the case of a water–ethanol mixture, the forces that exist when water and ethanol are mixed are stronger than water–water or ethanol–ethanol forces in the separate substances. Ethanol and water molecules can approach one another more closely in the mixture than either substance's molecules could approach a like molecule in the separate substances. Strong hydrogen bonding occurs in both ethanol and water.

23. (a) HF: Although both substances are capable of hydrogen bonding, water has two O—H bonds that can be involved in hydrogen bonding versus only one F—H bond in HF; (b) CH_3OCH_3: Because no H is attached to the O atom, no hydrogen bonding can exist. Thus, the molecule should be relatively more volatile than CH_3CH_2OH even though it contains the same number of atoms of each element; (c) CH_3SH: Hydrogen bonding is not as important for a SOH bond (because → has a lower electronegativity than O). Since there is relatively little hydrogen bonding, CH_3SH is more volatile than CH_3OH.

24. Both substances have the same molar mass. Ethyl alcohol contains a hydrogen atom directly bonded to an oxygen atom, however. Therefore, hydrogen bonding can exist in ethyl alcohol, whereas only weak dipole–dipole forces exist in dimethyl ether. Dimethyl ether is more volatile; ethyl alcohol has a higher boiling point.

26. *Ionic* solids have positive and negative ions as their fundamental particles; a simple example is sodium chloride, in which Na^+ and Cl^- ions are held together by strong electrostatic forces. *Molecular* solids have molecules as their fundamental particles, with the molecules being held together in the crystal by dipole–dipole forces, hydrogen-bonding forces, or London dispersion forces (depending on the identity of the substance); simple examples of molecular solids include ice (H_2O) and ordinary table sugar (sucrose). *Atomic* solids have simple atoms as their fundamental particles, with the atoms being held together in the crystal by either covalent bonding (as in graphite or diamond) or metallic bonding (as in copper or other metals).

28. Ionic solids consist of a crystal lattice of positively and negatively charged ions. A given ion is surrounded by several ions of the opposite charge, all of which electrostatically attract it strongly. This pattern repeats itself throughout the crystal. The existence of these strong electrostatic forces throughout the crystal means a great deal of energy must be applied to overcome the forces and melt the solid.

29. Ice contains nonlinear, highly polar water molecules, with extensive, strong hydrogen bonding. Dry ice consists of linear, nonpolar molecules, and only very weak intermolecular forces are possible.

30. Although ions exist in both the solid and liquid states, in the solid state the ions are rigidly held in place in the crystal lattice and cannot move so as to conduct an electric current.

제 14 장

2. A *non*homogeneous mixture may differ in composition in various places in the mixture, whereas a solution (a homogeneous mixture) has the same composition throughout. Examples of nonhomogeneous mixtures include spaghetti sauce, a jar of jelly beans, and a mixture of salt and sugar.

5. "Like dissolves like." The hydrocarbons in oil have intermolecular forces that are very different from those in water, so the oil spreads out rather than dissolving in the water.

8. unsaturated

10. large

12. 100.

13. (a) 0.0224%; (b) 18.3%; (c) 0.223%; (d) 18.3%

14. (a) 3.11 g NaCl, 121.89 g water; (b) 1.74 mg NaCl, 33.46 g water; (c) 62.1 g NaCl, 937.9 g water; (d) 0.0292 g NaCl, 29.17 g water

17. 7.81 g KBr

18. approximately 71 g

20. 0.110 mol; 0.220 mol

22. b

23. (a) 2.0 *M*; (b) 1.0 *M*; (c) 0.67 *M*; (d) 0.50 *M*

24. (a) 1.08 *M*; (b) 1.08 *M*; (c) 0.0108 *M*; (d) 0.108 *M*

25. 0.464 *M*

26. 0.0902 *M*

27. 0.619 *M*

28. (a) 0.0133 mol, 0.838 g; (b) 2.34 mol, 39.9 g; (c) 0.00505 mol, 0.490 g; (d) 0.0299 mol, 1.09 g

29. (a) 25.6 g; (b) 901 g; (c) 1.3 g; (d) 7.59 g

31. (a) 4.60×10^{-3} mol Al^{3+}, 1.38×10^{-2} mol Cl^-; (b) 1.70 mol Na^+, 0.568 mol PO_4^{3-}; (c) 2.19×10^{-3} mol Cu^{2+}, 4.38×10^{-3} mol Cl^-; (d) 3.96×10^{-5} mol Ca^{2+}, 7.91×10^{-5} mol OH^-
32. 1.33 g
34. half
36. (a) 0.0717 M; (b) 1.69 M; (c) 0.0426 M; (d) 0.625 M
38. 0.13 M
39. 10.3 mL
40. 31.2 mL
41. 0.523 g
42. 0.300 g
44. 1.8×10^{-4} M
45. (a) 63.0 mL; (b) 2.42 mL; (c) 50.1 mL; (d) 1.22 L
47. 1 *N*
49. (a) 0.277 *N*; (b) 3.37×10^{-3} *N*; (c) 1.63 *N*
50. (a) 0.134 *N*; (b) 0.0104 *N*; (c) 13.3 *N*
52. 22.2 mL, 11.1 mL
53. 0.05583 *M*, 0.1117 *N*

제 15 장

2. acid; base
4. A conjugate acid–base pair differs by one hydrogen ion, H^+. For example, $HC_2H_3O_2$ (acetic acid) differs from its conjugate base, $C_2H_3O_2^-$ (acetate ion), by a single H^+ ion.

$$HC_2H_3O_2(aq) \rightleftharpoons C_2H_3O_2\ (aq) + H\ (aq)$$

6. When an acid is dissolved in water, the hydronium ion (H_3O^+) is formed. The hydronium ion is the conjugate acid of water (H_2O).
7. (a) a conjugate pair (HSO_4^- is the acid, SO_4^{2-} is the base); (b) a conjugate pair (HBr is the acid, Br^- is the base; they differ by one proton); (c) not a conjugate pair ($H_2PO_4^{3-}$ is the conjugate acid of HPO_4^- and also the conjugate base of H_3PO_4; HPO_4^2 is the conjugate acid of PO_4^{3-}; (d) not a conjugate pair (NO_3^- is the conjugate base of HNO_3; NO_2^- is the conjugate base of HNO_2)
8. (a) $NH_3(aq)$(base) + $H_2O(l)$(acid) $\rightleftharpoons$ $NH_4^+(aq)$(acid) + $OH^-(aq)$(base);
(b) $NH_4^+(aq)$(acid) + $H_2O(l)$(base) $\rightleftharpoons$ $NH_3(aq)$(base) + $H_3O^+(aq)$(acid);
(c) $NH_4^-(aq)$(base) + $H_2O(l)$(acid) $\rightarrow$ $NH_3(aq)$(acid) + $OH^-(aq)$(base)
9. (a) $HBrO_3$; (b) HF; (c) HSO_3^-; (d) H_2SO_3
10. (a) BrO^-; (b) HSO_3^-; (c) SO_3^{2-}; (d) CH_3NH_2
11. (a) $CN^- + H_2O \rightleftharpoons HCN + OH^-$;
(b) $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$;
(c) $H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^{2-} + H_3O^+$;
(d) $NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$
13. If an acid is weak in aqueous solution, it does not easily transfer protons to water (and does not fully ionize).
If an acid does not lose protons easily, then the acid's anion must strongly attract protons.
14. A strong acid loses its protons easily and fully ionizes in water; the acid's conjugate base is poor at attracting and holding protons and is a relatively weak base. A weak acid resists loss of its protons and does not ionize to a great extent in water; the acid's conjugate base attracts and holds protons tightly and is a relatively strong base.
16. H_2SO_4 (sulfuric): $H_2SO_4 + H_2O \rightarrow HSO_4^- + H_3O^+$;
HCl (hydrochloric): $HCl + H_2O \rightarrow Cl^- + H_3O^+$;
HNO_3 (nitric): $HNO_3 + H_2O \rightarrow NO_3^- + H_3O^+$;
$HClO_4$ (perchloric): $HClO_4 + H_2O \rightarrow ClO_4^- + H_3O^+$
18. An oxyacid is an acid containing a particular element that is bonded to one or more oxygen atoms. HNO_3, H_2SO_4, and $HClO_4$ are oxyacids. HCl, HF, and HBr are not oxyacids.
19. (a) HSO_4 is a *moderately* strong acid; (b) HBr is a strong acid; (c) HCN is a weak acid; (d) $HC_2H_3O_2$ is a weak acid
21. HCO_3^- can behave as an acid if it reacts with a substance that more strongly gains protons than does HCO_3^- itself. For example, HCO_3^- would behave as an acid when reacting with hydroxide ion (a much stronger base): $HCO_3^-(aq) + OH^-(aq) \rightarrow CO_3^{2-}(aq) + H_2O(l)$. On the other hand, HCO_3^- would behave as a base when reacted with a substance that more readily loses protons than does HCO_3^- itself. For example, HCO_3^2 would behave as a base when reacting with hydrochloric acid (a much stronger acid): $HCO_3^-(aq) + HCl(aq) \rightarrow H_2CO_3(aq) + Cl^-(aq)$. $H_2PO_4^- + OH^- \rightarrow HPO_4^{2-} + H_2O$ and $H_2PO_4^- + H_3O^+ \rightarrow H_3PO_4 + H_2O$.
23. (a) $[H^+] = 2.4 \times 10^{-11}$ *M*, basic;
(b) $[H^+] = 1.3 \times 10^{-6}$ *M*, acidic;
(c) $[H^+] = 9.6 \times 10^{-13}$ *M*, basic;
(d) $[H^+] = 1.5 \times 10^{-8}$ *M*, basic
24. (a) 7.5×10^{-13} *M*, acidic; (b) 1.4×10^{-8} *M*, acidic;
(c) 2.5×10^{-6} *M*, basic; (d) 2.5×10^{-2} *M*, basic
25. (a) $[OH^-] = 0.105$ *M*; (b) $[OH^-] = 5.22 \times 10^{-5}$ *M*;
(c) $[OH\ ] = 8.41 \times 10^{-2}$ *M*
27. Answers will depend on student choices.
29. The pH of a solution is defined as the *negative* of the logarithm of the hydrogen ion concentration, pH $= -\log[H^+]$. Mathematically, the *negative* sign in the definition means the pH *decreases* as the hydrogen ion concentration *increases*.
30. (a) pH = 3.000 (acidic); (b) pH = 3.660 (acidic);
(c) pH = 10.037 (basic); (d) pH = 6.327 (acidic)
31. (a) 1.983, acidic; (b) 12.324, basic; (c) 11.368, basic;
(d) 3.989, acidic
32. (a) pOH = 6.55, basic; (b) pOH = 12.11, acidic;
(c) pOH = 0.85, basic; (d) pOH = 8.45, acidic
33. (a) pH = 1.719, $[OH^-] = 5.2 \times 10^{-13}$ *M*;
(b) pH = 6.316, $[OH^-] = 2.1 \times 10^{-8}$ *M*;
(c) pH = 10.050, $[OH^-] = 1.1 \times 10^{-4}$ *M*;
(d) pH = 4.212, $[OH^-] = 1.6 \times 10^{-10}$ *M*
34. (a) $[H^+] = 9.1 \times 10^{-10}$ *M*; (b) $[H^+] = 6.8 \times 10^{-6}$ *M*;
(c) $[H^+] = 5.0 \times 10^{-10}$ *M*; (d) $[H^+] = 2.6 \times 10^{-11}$ *M*
35. (a) $[OH^-] = 1.0 \times 10^{-13}$ *M*; (b) $[OH^-] = 9.8 \times 10^{-2}$ *M*;
(c) $[OH^-] = 6.0 \times 10^{-12}$ *M*; (d) $[OH^-] = 3.1 \times 10^{-2}$ *M*
36. (a) pH = 2.69; (b) pH = 9.86; (c) pH = 3.003;
(d) pH = 6.17
38. The solution contains water molecules, H_3O^+ ions (protons), and NO_3^- ions. Because HNO_3 is a strong acid that is completely ionized in water, no HNO_3 molecules are present.
39. (a) pH = 2.917; (b) pH = 3.701; (c) pH = 4.300;
(d) pH = 2.983
41. A buffered solution consists of a mixture of a weak acid and its conjugate base; one example of a buffered solution is a mixture of acetic acid ($HC_2H_3O_2$) and sodium acetate ($NaC_2H_3O_2$).
43. $CH_3COO^- + HCl \rightarrow CH_3COOH + Cl^-$;
$CH_3COOH + NaOH \rightarrow H_2O + NaCH_3COO$

제 16 장

1. The carbon–oxygen bonds in two carbon monoxide molecules and the oxygen–oxygen bond in an oxygen gas molecule must break, and the carbon–oxygen bonds in two carbon dioxide molecules must form.
2. E_a represents the *activation energy* for the reaction, which is the minimum energy needed for the reaction to be able to occur.
4. Enzymes are biochemical catalysts that speed up the com-

plicated reactions that would be too slow to sustain life at normal body temperatures.

5. A state of equilibrium is attained when two opposing processes are exactly balanced. The development of a vapor pressure above a liquid in a closed container is an example of a physical equilibrium. Any chemical reaction that appears to "stop" before completion is an example of a chemical equilibrium.
8. The two curves come together when a state of chemical equilibrium has been reached, after which point the forward and reverse reactions are occurring at the same rate so that there is no further net change in concentration.
10. (a) $K = \dfrac{[CO(g)]\,[H_2(g)]^3}{[CH_4(g)][H_2O(g)]}$; (b) $K = \dfrac{[O_3(g)]^2}{[O_2(g)]^3}$;
(c) $K = \dfrac{[CO_2(g)]^4[H_2O(g)]^6}{[C_2H_6(g)]^2[O_2(g)]^7}$;
11. (a) $K = \dfrac{[CH_3OH(g)]}{[CO(g)][H_2(g)]^2}$; (b) $K = \dfrac{[NO(g)]^2[O_2(g)]}{[NO_2(g)]^2}$;
(c) $K = \dfrac{[PBr_3(g)]^4}{[P_4(g)][Br_2(g)]^6}$
12. 9.4×10^{-20}
13. 1.2×10^{10}
15. (a) $K = \dfrac{[SF_6(g)]}{[F_2(g)]^3}$; (b) $K = \dfrac{[HCl(g)]^2}{[H_2S(g)][Cl_2(g)]}$;
(c) $K = \dfrac{[Cl_2O(g)]}{[SO_2(g)][Cl_2(g)]^2}$
16. (a) $K = \dfrac{[S_2Cl_2(g)]}{[CS_2(g)][Cl_2(g)]^3}$; (b) $K = \dfrac{1}{[Xe(g)][F_2(g)]^3}$;
(c) $K = \dfrac{1}{[O_2(g)]^3}$
17. $[CO_2]$ increases; K does not change
19. If heat is applied to an endothermic reaction (the temperature is raised), the equilibrium is shifted to the right. More product will be present at equilibrium than if the temperature had not been increased. The value of K increases.
20. (a) shift right; (b) no change (S is solid); (c) shift right; (d) no change
21. (a) no change (B is solid); (b) shift right; (c) shift left; (d) shift right
22. (a) no; (b) yes; (c) no; (d) yes
23. For an endothermic reaction, an increase in temperature will shift the position of equilibrium to the right (toward products).
25. A small equilibrium constant implies that not much product forms before equilibrium is reached. The reaction would not be a good source of the products unless Le Châtelier's principle can be used to force the reaction to the right.
26. 1.26×10^3
27. 1.06×10^{-1}
28. $[O_2(g)] = 8.0 \times 10^{-2}\,M$
29. $5.4 \times 10^{-4}\,M$
32. only the temperature
33. (a) $Ni(OH)_2(s) \rightleftharpoons Ni^{2+}(aq) + 2OH^-(aq)$, $K_{sp} = [Ni^{2+}(aq)][OH^-(aq)]^2$; (b) $Cr_2S_3(s) \rightleftharpoons 2Cr^{3+}(aq) + 3S^{2-}(aq)$, $K_{sp} = [Cr^{3+}(aq)]^2[S^{2-}(aq)]^3$; (c) $Hg(OH)_2(s) \rightleftharpoons Hg^{2+}(aq) + 2OH^-(aq)$, $K_{sp} = [Hg^{2+}(aq)][OH^-(aq)]^2$; (d) $Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$, $K_{sp} = [Ag^+(aq)]^2[CO_3^{2-}(aq)]$
34. $6.5 \times 10^{-5}\,M$; 0.021 g/L
35. 7.4×10^{-4} g/L
36. $K_{sp} = 2.3 \times 10^{-47}$
37. $K_{sp} = 1.23 \times 10^{-15}$
38. $K_{sp} = 1.9 \times 10^{-4}$; 10. g/L
70. $4 \times 10^{-17}\,M$, 4×10^{-15} g/L

제 17 장

2. Oxidation is a loss of one or more electrons by an atom or ion. Reduction is the gaining of one or more electrons by an atom or ion. Equations depend on student responses.
3. (a) Potassium is oxidized, oxygen is reduced; (b) iodine is oxidized, chlorine is reduced; (c) cobalt is oxidized, chlorine is reduced; (d) carbon is oxidized, oxygen is reduced
4. (a) sulfur is oxidized, oxygen is reduced;
(b) phosphorus is oxidized, oxygen is reduced;
(c) hydrogen is oxidized, carbon is reduced;
(d) boron is oxidized, hydrogen is reduced
6. Oxidation numbers represent a "relative charge" one atom has compared to another in a compound. In an element, all the atoms are equivalent.
8. Because fluorine is the most electronegative element, its oxidation state is always negative relative to other elements; because fluorine gains only one electron to complete its outermost shell, its oxidation number in compounds is always -1. The other halogen elements are almost always more electronegative than the atoms to which they bond, and almost always have -1 oxidation numbers. However, in an interhalogen compound involving fluorine and some other halogen, since fluorine is the most electronegative element of all, the other halogens in the compound will have posi- tive oxidation states relative to fluorine.
10. 3–
11. (a) Cr, +3; Cl, −1; (b) Cu, +1; O, −2;
(c) Cu, +2; O, −2; (d) 0
12. (a) Al, +3; P, +5; O, −2; (b) Mn, +4; O, −2;
(c) Ba, +2; C, +4; O, −2; (d) Cl, +1; F, −1
13. (a) H, +1; P, +5; O, −2; (b) H, +1; Br, +1; O, −2;
(c) H, +1; N, +5; O, −2; (d) H, +1; Cl, +7; O, −2
14. (a) K, +1; Cl, +5; O, −2; (b) 0; (c) C, +2; O, −2;
(d) Na, +1; I, +5; O, −2
15. (a) Fe, +3; O, −2; (b) Al, +3; C, +4; O, −2;
(c) Ba, +2; Cr, +6; O, −2; (d) Ca, +2; H, +1; C, +4; O, −2
17. Electrons are negative; when an atom gains electrons, it gains one negative charge for each electron gained. For example, in the reduction reaction $Cl + e^- \rightarrow Cl^-$, the oxidation state of chlorine decreases from 0 to -1 as the electron is gained.
19. An oxidizing agent oxidizes another species by gaining the electrons lost by the other species; therefore, an oxidizing agent itself decreases in oxidation state. A reducing agent increases its oxidation state when acting on another atom or molecule.
21. (a) manganese is oxidized, hydrogen is reduced;
(b) sulfur is oxidized, oxygen is reduced;
(c) aluminum is oxidized, hydrogen is reduced;
(d) nitrogen is oxidized, oxygen is reduced
22. (a) carbon is oxidized, chlorine is reduced;
(b) carbon is oxidized, oxygen is reduced;
(c) phosphorus is oxidized, chlorine is reduced;
(d) calcium is oxidized, hydrogen is reduced
23. Iron is reduced [+3 in $Fe_2O_3(s)$, 0 in $Fe(l)$]; carbon is oxidized [+2 in $CO(g)$, +4 in $CO_2(g)$]. $Fe_2O_3(s)$ is the oxidizing agent; $CO(g)$ is the reducing agent.
24. (a) chlorine is reduced, iodine is oxidized; chlorine is the oxidizing agent, iodide ion is the reducing agent; (b) iron is reduced, iodine is oxidized; iron(III) is the oxidizing agent, iodide ion is the reducing agent; (c) copper is reduced, iodine is oxidized; copper(II) is the oxidizing agent, iodide ion is the reducing agent

26. Under ordinary conditions it is impossible to have "free" electrons that are not part of some atom, ion, or molecule. Thus, the total number of electrons lost by the species being oxidized must equal the total number of electrons gained by the species being reduced.

27. (a) $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$; (b) $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$; (c) $Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$; (d) $Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$

28. (a) $8e^- + 10H^+(aq) + NO_3^-(aq) \rightarrow NH_4^+(aq) + 3H_2O(l)$;
(b) $2e^- + 2H^+(aq) + C_2N_2(g) \rightarrow 2HCN(aq)$;
(c) $6e^- + 6H^+(aq) + ClO_3^-(aq) \rightarrow Cl^-(aq) + 3H_2O(l)$;
(d) $2e^- + 4H^+(aq) + MnO_2(s) \rightarrow Mn^{2+}(aq) + 2H_2O(l)$

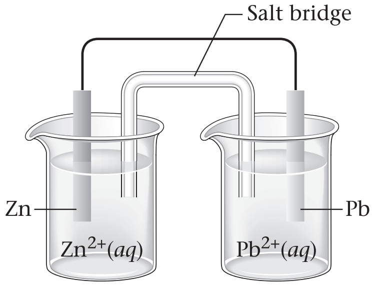
29. (a) $2Al + 6H^+ \rightarrow 2Al^{3+} + 3H_2$; (b) $8H^+ + 2NO_3^- + 3S^{2-} \rightarrow 3S + 2NO + 4H_2O$; (c) $6H_2O + I_2 + 5Cl_2 \rightarrow 2IO_3^- + 2H^+ + 10HCl$; (d) $2H^+ + AsO_4^- + S^{2-} \rightarrow + AsO_3^- + H_2O$

30. $Cu(s) + 2HNO_3(aq) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$; $Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$

31. A salt bridge typically consists of a U-shaped tube filled with an inert electrolyte (one involving ions that are not part of the oxidation–reduction reaction). A salt bridge completes the electrical circuit in a cell. Any method that allows transfer of charge without allowing bulk mixing of the solutions may be used (another common method is to set up one half-cell in a porous cup, which is then placed in the beaker containing the second half-cell).

33. Reduction takes place at the cathode and oxidation takes place at the anode.

34.



$Pb^{2+}(aq)$ ion is reduced; $Zn(s)$ is oxidized. The anode reaction is $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$. The cathode reaction is $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$.

36. $Cd + 2OH^- \rightarrow Cd(OH)_2 + 2e^-$ (oxidation);
$NiO_2 + 2H_2O + 2e^- \rightarrow Ni(OH)_2 + 2OH^-$ (reduction)

38. Aluminum is a very reactive metal when freshly isolated in the pure state. Upon standing for even a relatively short period of time, aluminum metal forms a thin coating of Al_2O_3 on its surface from reaction with atmospheric oxygen. This Al_2O_3 coating is much less reactive than the metal and protects the metal's surface from further attack.

40. The magnesium is used for cathodic protection of the steel pipeline. Magnesium is more reactive than iron, and will be oxidized in preference to the iron of the pipeline.

42. The main recharging reaction for the lead storage battery is $2PbSO_4(s) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(s) + 2H_2SO_4(aq)$. A major side reaction is the electrolysis of water, $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$, which produces an explosive mixture of hydrogen and oxygen that accounts for many accidents during the recharging of such batteries.

제 18 장

2. The radius of a typical atomic nucleus is on the order of 10^{-13} cm, which is roughly 100,000 times smaller than the radius of an atom overall.

4. The mass number represents the total number of protons and neutrons in a nucleus.

5. The atomic number (Z) is written as a left subscript, while the mass number (A) is written as a left superscript. That is, the general symbol for a nuclide is ${}^{A}_{Z}X$. As an example, consider the isotope of oxygen with 8 protons and 8 neutrons: its symbol would be ${}^{16}_{8}O$.

7. a neutron

10. Gamma rays are high-energy photons of electromagnetic radiation; they are not normally considered to be particles. When a nucleus produces only gamma radiation, the atomic number and mass number of the nucleus do not change.

12. Electron capture occurs when one of the inner-orbital electrons is pulled into, and becomes part of, the nucleus.

14. ${}^{7}_{3}Li$ and ${}^{6}_{3}Li$; since there is a much larger abundance of ${}^{7}_{3}Li$, the average mass number will be closer to 7.

16. (a) electron; (b) positron; (c) neutron; (d) proton

17. (a) ${}^{0}_{-1}e$; (b) ${}^{14}_{7}N$; (c) ${}^{0}_{-1}e$

18. (a) ${}^{218}_{86}Rn$; (b) ${}^{0}_{1}e$ (positron); (c)

19. (a) ${}^{137}_{55}Cs \rightarrow {}^{0}_{-1}e + {}^{137}_{56}Ba$; (b) ${}^{3}_{1}H \rightarrow {}^{0}_{-1}e + {}^{3}_{2}He$; (c) ${}^{216}_{84}Po \rightarrow {}^{0}_{-1}e + {}^{216}_{85}At$

23. ${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{15}P + {}^{1}_{0}n$

24. The half-life of a nucleus is the time required for one-half of the original sample of nuclei to decay. A given isotope of an element always has the same half-life, although different isotopes of the same element may have greatly different half-lives. Nuclei of different elements have different half-lives.

25. ${}^{226}_{88}Ra$ is the most stable (longest half-life); ${}^{224}_{88}Ra$ is the "hottest" (shortest half-life).

28. After four half-lives, a little over 6 μg

29. For an administered dose of 100 μg, 0.39 μg remains after 2 days. The fraction remaining is 0.39/100 = 0.0039; on a percentage basis, less than 0.4% of the original radioisotope remains.

31. Carbon-14 is produced in the upper atmosphere by the bombardment of nitrogen with neutrons from space:

$$ {}^{14}_{7}N + {}^{1}_{0}n \rightarrow {}^{14}_{6}C + {}^{1}_{1}H $$

33. We assume that the concentration of C-14 in the atmosphere is effectively constant. A living organism is constantly replenishing C-14 through the processes of either metabolism (sugars ingested in foods contain C-14) or photosynthesis (carbon dioxide contains C-14). When a plant dies, it no longer replenishes itself with C-14 from the atmosphere. As the C-14 undergoes radioactive decay, its amount decreases with time.

35. (a) thyroid gland; (b) heart muscle; (c) bones, heart, liver, lungs; (d) circulatory system

37. fission, fusion, fusion, fission

39. ${}^{1}_{0}n + {}^{235}_{92}U \rightarrow {}^{142}_{56}Ba + {}^{91}_{36}Kr + 3{}^{1}_{0}n$

41. A critical mass of a fissionable material is the amount needed to provide a high enough internal neutron flux to sustain the chain reaction (production of enough neutrons to cause the continuous fission of further material). A sam- ple with less than a critical mass is still radioactive, but can- not sustain a chain reaction.

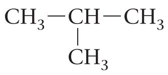
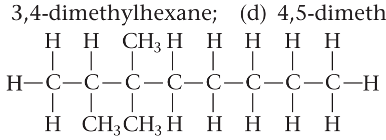
43. An actual nuclear explosion, of the type produced by a nuclear weapon, cannot occur in a nuclear reactor because the concentration of the fissionable materials is not sufficient to form a supercritical mass.

46. protons (hydrogen), helium

47. Somatic damage is damage directly to the organism itself, causing nearly immediate sickness or death to the organism. Genetic damage is damage to the genetic machinery of the organism, which will be manifested in future generations of offspring.

49. Gamma rays penetrate long distances, but seldom cause ionization of biological molecules. Because they are much heavier, although less penetrating, alpha particles ionize biological molecules very effectively and leave a dense trail of damage in the organism. Isotopes that decay by releasing alpha particles can be ingested or breathed into the body, where the damage from the alpha particles will be more acute.

제 19 장

2. A given carbon atom can be attached to a maximum of four other atoms. Carbon atoms have four valence electrons. By making four bonds, carbon atoms exactly complete their valence octet.
4. A triple bond represents the sharing of six electrons (three pairs of electrons). The simplest example of an organic molecule containing a triple bond is acetylene,
H : C : : : C : H (H—C≡C—H)
6. $\ddot{O}=C=\ddot{O}$; $C\equiv O$
7. (a) $CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$;
(b) $CH_3—CH_2—CH_2—CH_2—CH_3$;
(c) $CH_3—CH_2—CH_3$;
(d) $CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$
8. (a) $CH_3—CH_2—CH_2—CH_2—CH_2—CH_3$;
(b) $CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$;
(c) $CH_3—CH_2—CH_2—CH_3$;
(d) $CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$
10. A branched alkane contains one or more shorter carbon-atom chains attached to the side of the main (longest) carbon-atom chain. The simplest branched alkane is 2-methylpropane,

12. Structures depend on student choices.
15. Multiple substituents are listed in alphabetical order, disregarding any prefix.
16. (a) 2,3,4-trimethylpentane; (b) 2,3-dimethylpentane; (c) 3,4-dimethylhexane; (d) 4,5-dimethyloctane
17. (a)

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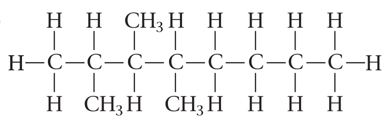
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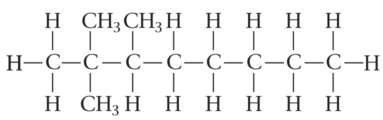
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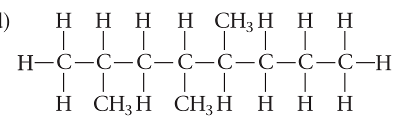
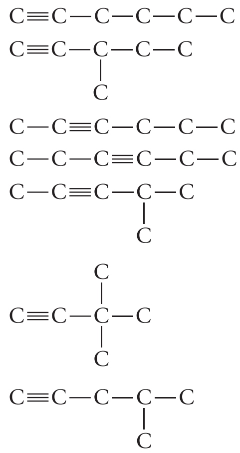
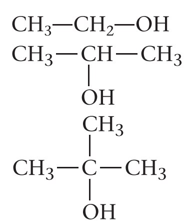
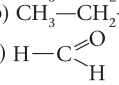
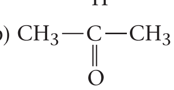
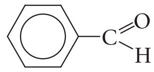

18.

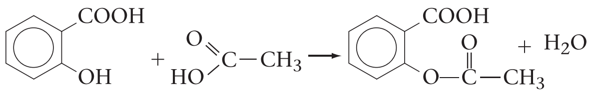
Number of C Atoms	Use
$C_5–C_{12}$	gasoline
$C_{10}–C_{18}$	kerosene, jet fuel
$C_{15}–C_{25}$	diesel fuel, heating oil
$>C_{25}$	asphalt

19. Tetraethyl lead was added to gasoline to prevent "knocking" of high-efficiency automobile engines. Its use is being discontinued because of the danger to the environment from the lead in this substance.
20. Combustion reactions are used as a source of heat and light:
$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
21. (a) $2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$;
(b) $CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(l) + HCl(g)$;
(c) $CHCl_3(l) + Cl_2(g) \rightarrow CCl_4(l) + HCl(g)$
23. An alkyne is a hydrocarbon containing a carbon – carbon triple bond. The general formula is C_nH_{2n-2}.
25. (a) 2-chloro-1-butene; (b) 3-chloro-1-butene; (c) 3-chloro-2-butene; (d) 1-chloro-2-butene
26. The carbon skeletons are

27. For benzene, a *set* of equivalent Lewis structures can be drawn, with each structure differing only in the location of the three double bonds in the ring. Experimentally, benzene does not demonstrate the chemical properties expected for molecules having *any* double bonds.
29. *ortho*-: adjacent substituents (1,2-); *meta*-: two substituents with one unsubstituted carbon atom between them (1,3-); *para*-: two substituents with two unsubstituted carbon atoms between them (1,4-)
30. (a) anthracene; (b) 1,3,5-trimethylbenzene; (c) 1,4-dinitrobenzene; (d) 4-bromotoluene (4-bromo-1-methylbenzene)
32. (a) organic acids; (b) aldehydes; (c) ketones; (d) ethers
33. Primary alcohols have one hydrocarbon fragment (alkyl group) bonded to the carbon atom where the —OH group is attached. Secondary alcohols have two alkyl groups attached, and tertiary alcohols contain three alkyl groups. Examples are
ethanol (primary)
2-propanol (secondary)
2-methyl-2-propanol (tertiary)

35. $C_6H_{12}O_6 \xrightarrow{\text{yeast}} 2CH_3—CH_2—OH + 2CO_2$
The yeast necessary for the fermentation process are killed if the concentration of ethanol is greater than 13%. More concentrated ethanol solutions are most commonly made by distillation.
36. Methanol (CH_3OH): starting material for synthesis of acetic acid and many plastics; ethylene glycol ($CH_2OH—CH_2OH$): automobile antifreeze; isopropyl alcohol (2-propanol, $CH_3—CH(OH)—CH_3$): rubbing alcohol
38. (a) $CH_3—CH_2—CH_2—COOH$;
(b) $CH_3—CH_2—C(=O)—CH_3$
40. (a)

(b)


(c) $CH_3-\underset{\displaystyle \overset{\|}{O}}{C}-CH_2-CH_3$

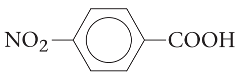
(d) 

42. (a) $CH_3-CH_2-CH_2-CHO$;
(b) CH_3-CH_2-COOH;
(c) $CH_3-CH_2-CH_2-C(=O)-O-CH_2-CH_2-CH_3$

80. 

82. (a) $CH_3-CH_2-CH_2-\underset{\displaystyle CH_3}{\underset{|}{CH}}-\underset{\displaystyle CH_3}{\underset{|}{CH}}-COOH$

(b) $CH_3-CH_2-C(=O)-O-CH_2-CH_3$

(c) 

(d) $CH_3-C(=O)-O-CH(CH_3)_2$

48. In addition polymerization, the monomer units add together to form the polymer, with no other products. Polyethylene and polytetrafluoroethylene (Teflon) are examples.

50. A polyester is formed from the reaction of a dialcohol (two —OH groups) with a diacid (two —COOH groups). One —OH group of the alcohol forms an *ester linkage* with one of the —COOH groups of the acid. The resulting dimer possesses an —OH and a —COOH group, so the dimer can undergo further esterification reactions. Dacron is an example.