Answer Key

Unit 3 Quantities in Chemical Reactions

Unit Preparation Questions (Assessing Readiness)

(Student textbook page 218–22)

1. d

2. d

3. Determine whether the compound contains any metal atoms. A compound with both metal and non-metal atoms will likely be ionic. A compound with only non-metal atoms will likely be molecular.

4. calcium: 3, phosphorus: 2; oxygen: 8

5. a. ionic, LiI
b. molecular, CF_4
c. ionic, SnS_2
d. molecular, N_2O_5
e. ionic, AlH_3
f. molecular, ClF_3

6. b

7. a. III **b.** V **c.** II **d.** IV **e.** I

8. a. colour change, precipitate formed
b. heat and light are produced

9. a. two or more elements combine to form a single compound
b. a compound and an element form a different element and a different compound
c. one compound splits into two or more elements and/or compounds
d. two compounds change into two different compounds
e. a hydrocarbon fuel reacts completely with oxygen to form carbon dioxide and water
f. a hydrocarbon fuel reacts partially with oxygen to form a mixture of carbon dioxide, water, carbon monoxide, and soot (pure carbon)

10. a. single displacement
b. synthesis
c. decomposition
d. synthesis
e. double displacement
f. decomposition

11. b

12. d

13. From the law of conservation of mass, we know that atoms are neither created nor destroyed in a chemical reaction. By balancing a chemical equation, we show exactly what happened to each atom. The balanced ratios allow us to assess relative quantities of reactants used and products formed in the reaction.

14. The sulfate was tripled in the sulfuric acid (without tripling the hydrogen), turning it into a different substance (rather than a larger quantity of sulfuric acid).

15. $3H_2SO_4(aq) + 2Al(s) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

16. a. $Br_2(\ell) + 2NaI(aq) \rightarrow 2NaBr(aq) + I_2(s)$
b. $2Al(s) + 3Cu(NO_3)_2(aq) \rightarrow 3Cu(s) + 2Al(NO_3)_3(aq)$
c. $2Fe_2O_3(s) \rightarrow 4Fe(s) + 3O_2(g)$
d. $C_5H_{12}(\ell) + 8O_2(g) \rightarrow 6H_2O(g) + 5CO_2(g)$
e. $6\ Li(s) + N_2(g) \rightarrow 2\ Li_3N(s)$
f. $2AgNO_3(aq) + CaCl_2(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$

17. a. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$; synthesis
b. $CaCl_2(aq) + 2NH_4OH(aq) \rightarrow Ca(OH)_2(s) + 2NH_4Cl(aq)$; double displacement
c. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$; combustion

18. a. $8Zn(s) + S_8(s) \rightarrow 8ZnS(s)$; solid formed and colour changed
b. $H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(\ell)$; bubbles (gas formed)
c. $BaCl_2(aq) + K_2CrO_3(aq) \rightarrow BaCrO_3(s) + 2KCl(aq)$; precipitate formed
d. $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$; energy change (light and heat)

19. a. $2KClO_3(s) \rightarrow 2\ KCl(s) + 3O_2(g)$
potassium chlorate à potassium chloride + oxygen
decomposition
b. $2Al(s) + 3H_2SO_4(aq) \rightarrow 3H_2(g) + Al_2(SO_4)_3(aq)$
aluminum + sulfuric acid → hydrogen + aluminum sulfate
single displacement

c. $3Ba(OH)_2(aq) + 2H_3PO_4(aq) \rightarrow 6H_2O(\ell) + Ba_3(PO_4)_2(s)$
barium hydroxide + phosphoric acid →
water + barium phosphate
double displacement (neutralization)

20. c

21. b

22. Yes. The product of reaction e would cause litmus paper to turn red, indicating an acid. The product of reaction b might also be acidic if the reaction took place in water (dissolved CO_2 is acidic). The product of reaction c would cause litmus paper to turn blue, indicating a base.

23. Neutralization is a reaction in which an acid and a base react to form a salt and water.

24. The hydrogen ion (H^+) or, in water, the hydronium ion ($H_3O^+(aq)$), are characteristic of acids. The hydroxide ion, OH^- is characteristic of bases. Hydrogen/hydronium ions have a positive charge, and hydroxide ions have a negative charge.

25. By testing a neutralization reaction with litmus paper, you can tell when the reaction is complete when the litmus paper does not change colour (indicating a neutral substance).

26. a. low precision, low accuracy
b. high precision, low accuracy
c. high precision, high accuracy

27. a. 3 **b.** 3 **c.** 2 **d.** 4 **e.** 1

28. a. 1.86 kg **b.** 0.4 L **c.** 0.65 m
d. 3.269×10^4 kg **e.** 15.2 J **f.** 0.0080 km

29. b, d, e

30. a. 2.4×10^2 **d.** 1.6×10^6
b. 5.3×10^{-3} **e.** 2.0×10^{-10}
c. 4.2×100

31. 2.18 kg

32. a. 8.41 m **b.** 52 g/mL **c.** 16.1 L
d. 0.28 km **e.** 17.22 kg **f.** 9.05 Mg

33. a. 6.9 mg **d.** 1.4×10^2 kg • m/s^2
b. 22.5 m/s^2 **e.** 0.88 m/s
c. 8.3 kPa • L/mol • K

34. a. 81% **b.** \$11.19, \$12.64

35. a. 35 **b.** 2 **c.** 0.4 **d.** 12 **e.** 165 **f.** 7

Chapter 5 The Mole: A Chemist's Counter

Learning Check Questions

(Student textbook page 226)

1. The Avogadro constant is defined as exactly equal to the number of atoms of carbon-12 in 1 g of carbon-12. However, the numerical value must be determined experimentally. Scientists are constantly updating the value as they improve the methods used to measure the number of atoms in 1 g of carbon-12.

2. A mole is a unit used to measure the amount of a substance. One mole contains the same number of particles as the number of atoms in 1 g of carbon-12.

3. You would have two times the Avogadro constant of hydrogen atoms. Rounded off, the number would be $2(6.02 \times 10^{23}) = 1.20 \times 10^{24}$ hydrogen atoms.

4. You would not be able to see one person, but a mole of people is so many that they would be visible, as a group, from space. In fact, a mole of people would have a mass about the same size as the mass of Earth.

5. four

6. Paper, like slices of bread, is used in larger quantities. It is much more efficient to count these items in larger units, such as reams and loaves.

7. Oxygen was originally chosen by chemists as the reference for measuring atomic masses. However, once the mass spectrometer was developed as a way to measure atomic masses, the reference was changed to carbon. Mass spectrometers accelerate particles in a vacuum, through a magnetic field. To create a vacuum, strong pumps must be used. A tiny amount of the carbon from the pump lubricants always got into the vacuum. Most naturally-occurring carbon is carbon-12. So, since it was already there, chemists decided to use carbon-12 as the universal reference for atomic masses.

(Student textbook page 234)

8. Since we do not have a device that can count the number of particles in a substance directly, it is not convenient to measure substances in moles. Instead, we can measure substances in grams, and then convert to moles using a simple calculation.

9. Atomic molar mass refers to the mass of one mole of atoms. Molar mass is a more general term that refers to the mass of one mole of any type of particle, including atoms, molecules, and formula units.

10. The numerical values for the atomic mass of an element in atomic mass units (u) and the atomic molar mass of the same element in grams (g) are the same.

11. 63.55 g

12. The mass of 6.02×10^{23} atoms of a particular element can also be described as the atomic molar mass (or just molar mass) of that element. The units are g/mol.

13. Molar masses are usually not whole numbers because the reported values of atomic masses are weighted averages of naturally occurring isotopes. In addition, the mass of an individual isotope as measured by a mass spectrometer is not a whole number.

Caption Questions

Figure 5.1 (Student textbook page 224): For example, a score is 20, a gross is 12×12 or 144, a great gross is twelve gross or 1728, a baker's dozen is 13, and a paper bale is 5000 sheets of paper.

Figure 5.5 (Student textbook page 230): 6.02×10^{22} molecules

Figure 5.10 (Student textbook page 236): 0.496 mol

Section 5.1 Review Questions

(Student textbook page 232)

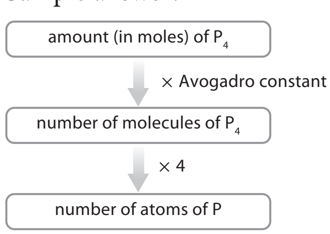
1. One mole of a substance contains 6.02×10^{23} atoms, molecules, or formula units of that substance.
For example, one mole of helium atoms contains 6.02×10^{23} helium atoms.

2. A mole and a dozen are similar in that both represent a specific number of things. They are different in that a mole represents a much bigger number, 6.02×10^{23}, and is used to deal with atoms, molecules, ions, and formula units. A dozen only represents the number 12, and is used to deal with macroscopic items.

3. No, mole value is not exact, because methods are still improving, so the number of atoms in 1 g of carbon-12 continues to be revised (made more precise).

4. 2.993×10^{-26} L

5. Sample answer:



6. 3.48×10^{22} formula units

7. a. 9.39×10^{22} atoms
b. 4.7×10^{24} formula units
c. 9.15×10^{24} molecules

8. a. 4.5×10^{-4} mol
b. 0.0117 mol

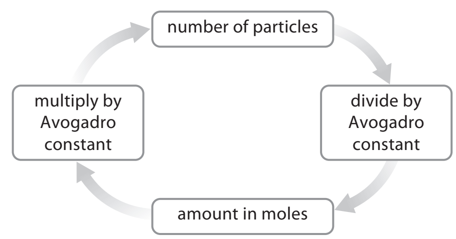
9. The sample of carbon

10. 2.37×10^{-3} mol

11. a. 1.6×10^{24} formula units
b. 7.8×10^{24} atoms
c. 3.1×10^{24} atoms

12. a. 2.8×10^{23} formula units
b. 5.5×10^{23} atoms

13. Sample answer:



14. octane, sodium hydrogen carbonate, copper

15. 7.224×10^{24} atoms

Section 5.2 Review Questions

(Student textbook page 243)

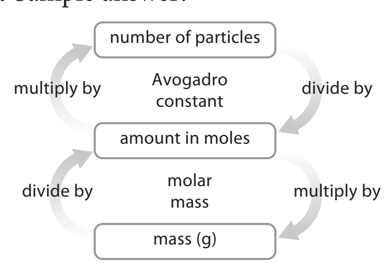
1. a. Atomic molar mass refers to the molar mass of atoms, while molar mass is a more general term referring to the molar mass of any type of particle, including atoms, molecules, and formula units.

b. Both terms refer to a specific type of molar mass, but molecular molar mass refers to the molar mass of molecules, while formula unit molar mass refers to the molar mass of formula units.

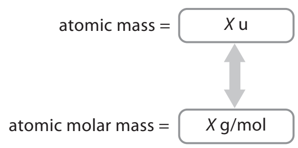
c. Both terms refer to a specific type of molar mass, but formula unit molar mass refers to the molar mass of formula units, while atomic molar mass refers to the molar mass of atoms.

d. Molar mass is a more general term, referring to any type of particle. Molecular molar mass refers to the molar mass of molecules.

2. Sample answer:



3. Sample answer:



4. 0.241 mol
5. 1.0×10^5 g
6. 1.80×10^2 mol
7. 2.7×10^2 g
8. 5.2×10^4 g
9. 2.42×10^{23} atoms
10. the aluminum sulfate sample
11. the carbon dioxide sample
12. **a.** 2.1×10^{17} formula units
    b. 6.3×10^{17} atoms
13. 3.1×10^{17} atoms of lead in 1 L translates into 0.11 mg/L, which is about ten times greater than the allowable lead limit. Therefore, the water is not safe to drink.
14.

Substance	Number of Particles	Amount (mol)	Mass (g)
$P_4(s)$	7.95×10^{24}	13.2	1.64×10^3
$Ba(MnO_4)_2(s)$	6.7×10^{20}	1.1×10^{-3}	0.42
$C_5H_9NO_{4(s)}$	8.027×10^{22}	0.1333	19.62

15. 3.49×10^{-4} g
16. acetic acid, oxalic acid, benzoic acid

Practice Problems

(Student textbook page 228)

1. 9×10^{13} refrigerators
2. 1.6×10^{27} km
3. 1.91×10^{16} years
4. 5.8×10^{16} km high
5. 2.1×10^{11} Rogers Centres
6. 2.48×10^{15} rows
7. One mole of tablespoons has a volume of 9.03×10^9 km^3. Because this volume is greater than the total volume of the oceans, you would drain the oceans. In fact, you could drain the equivalent of over *six times* the world's oceans.
8. 1.91×10^{12} dollars/s
9. Earth is 4.1×10^3 times heavier.
10. 1.2×10^{16} cm

(Student textbook page 230)

11. 6.38×10^{22} atoms
12. 5×10^{21} atoms
13. 5.1×10^{27} molecules
14. 5.11×10^{26} formula units
15. 3.15×10^{22} formula units
16. 2.32×10^{24} molecules
17. 1.3×10^{27} atoms
18. **a.** 2.90×10^{24} molecules **b.** 1.45×10^{25} atoms
19. **a.** 4.36×10^{25} atoms **b.** 2.18×10^{26} molecules
20. **a.** 0.015 mol **b.** 1.05×10^{23} atoms

(Student textbook page 231)

21. 0.158 mol
22. 0.277 mol
23. 2.0×10^2 mol
24. 1.4×10^{-4} mol
25. 5.1×10^4 mol
26. 27.5 mol
27. 2.0 mol
28. 0.0346 mol
29. 1.3×10^3 mol
30. 0.106 mol

(Student textbook page 235)

31. **a.** 22.99 g/mol **b.** 183.84 g/mol **c.** 131.29 g/mol **d.** 58.69 g/mol
32. 123.88 g/mol
33. 310.18 g/mol
34. 331.2 g/mol
35. 113.94 g/mol
36. 306.52 g/mol
37. 315.51 g/mol
38. 283.88 g/mol
39. 392.21 g/mol
40. 132.91 g/mol

(Student textbook page 237)

41. 182 g
42. 11 g
43. 0.231 g or 231 mg

44. 5.3×10^2 mg

45. a. cobalt(II) nitrate, 8.2×10^{-1} g

b. lead(IV) thiosulfate, 1.28×10^4 g

46. a. NH_4NO_3, 3.9×10^2 g **b.** Fe_2O_3, 2.59×10^3 g

47. 2.4×10^2 mg

48. 1.001 kg

49. a. bromine **b.** strontium iodate

50. aluminum iodate

(Student textbook page 239)

51. 1.73 mol

52. 139 mol

53. 8.75×10^{-4} mol

54. 1.1×10^{-4} mol

55. a. SiO_2, 6.2×10^{-5} mol

b. $Ti(NO_3)_4$, 0.08577 mol

c. $In_2(CO_3)_3$, 4.69×10^{-5} mol

d. $CuSO_4 \bullet 5H_2O$, 313 mol

56. 1.47 mol

57. 1.80×10^2 mol

58. 1.52×10^{-5} mol

59. $Al(OH)_3(s)$, $AgCl(s)$, $Ni(NO_3)_2(s)$

60. barium perchlorate < glucose < tin(IV) oxide

(Student textbook page 242)

61. a. 3.52×10^3 g **b.** 616 g **c.** 7.00 g **d.** 1.8×10^6 g

62. a. 2.40×10^{22} formula units

b. 4.82×10^{19} molecules

c. 2.0×10^{26} formula units

d. 1.75×10^{24} formula units

63. a. 9.52 g

b. 2.51×10^{24} formula units

64. a. 2.1×10^{18} molecules

b. 1.1 mg/day

65. 3.24×10^{-22} g

66. a. 2.58×10^{24} atoms

b. 6.88×10^{24} atoms

c. 8.60×10^{23} atoms

67. a. 1.26×10^{22} formula units

b. 6.29×10^{22} ions

68. gallium arsenide

69. 0.37 g

70. hydrogen cyanide, acetic acid, sucrose

Chapter 5 Review Questions

(Student textbook pages 251–3)

1. c

2. a

3. b

4. a

5. d

6. c

7. c

8. b

9. No, there are three atoms in each molecule of water, so there are three times as many atoms as there are molecules.

10. 65.41; 6.02×10^{23}

11. a. One mole of chromium and one mole of aluminum both have the same number of particles: 6.02×10^{23}.

b. The mass of one mole of each element is different, because the atoms are different sizes. One mole of chromium has a mass of 52.00 g, while one mole of aluminum has a mass of 36.98 g.

12. First, determine the atomic molar mass of each element in the compound. Next, multiply each value by the number of atoms of each element in one molecule. Finally, add all the calculated values together to obtain the molar mass.

13. First, determine the molar mass of the compound. Next, multiply the amount of moles by the molar mass to obtain the mass.

14. The Avogadro constant, 6.02×10^{23}

15. The value 65.41 can have atomic mass units (u), which describe the average mass of one atom of zinc. This value can also have the units of grams per mole (g/mol), which describe the mass in grams of one mole of zinc atoms.

16. Since the mole is a counting unit, it can be used to describe the number of chemical atoms, molecules, ions, and formula units that react in an equation. However, since there is no device that can count the amount of moles in a sample of a substance, we work in mass, a measurable unit, to measure out substances.

17. a. 0.36 mol **b.** 15.8 mol **c.** 5.91×10^{-22} mol

18. a. 1.2 g **b.** 119 g **c.** 1.4×10^5 g

19. 1.94×10^{-3} mol

20. 1.2×10^4 g

21. 20.0 g glucose = 0.111 mol; 20.0 g propane = 0.453 mol

22. nitrogen trichloride

23. a. 0.06802 mol **b.** 0.08752 mol **c.** 0.09331 mol

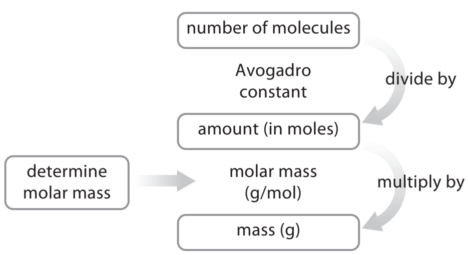
24. We use atomic mass units (u) to record the mass of individual atoms or other particles. We use atomic molar mass to record the mass in grams of one mole of atoms (g/mol). Atomic molar mass is more useful to a chemist because it refers to a more practical amount (moles of atoms rather than individual atoms) and because mass in grams can be measured in the lab.

25.

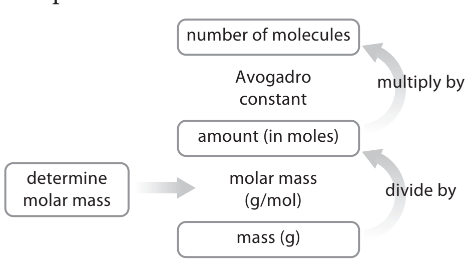
Substance	Total Number of Atoms	Number of Molecules or Formula Units	Molar Mass (g/mol)	Amount of Substance (mol)	Mass (g)
$C_3H_6O_2(\ell)$	4.47×10^{23}	4.06×10^{22}	74.09	0.0675	5.00
$NaC_6H_5COO(s)$	3.56×10^{19}	2.37×10^{18}	144.11	3.94×10^{-6}	5.68×10^{-4}
$Al(H_2PO_4)_3(s)$	1.363×10^{24}	6.193×10^{22}	317.95	0.1029	32.71
$CCl_2F_2(g)$	2.38×10^{26}	4.75×10^{25}	120.9	78.9	9.54×10^{3}
$C_4H_{10}O_2(\ell)$	5.53×10^{23}	3.46×10^{22}	90.12	0.0574	5.17
$NaHCO_3(s)$	1.778×10^{24}	2.963×10^{23}	84.01	0.4921	41.34

26. $C_{11}H_{12}N_2O_2$

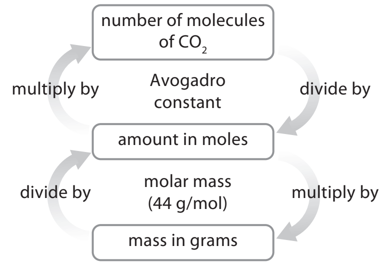
27. Sample answer:



28. Sample answer:

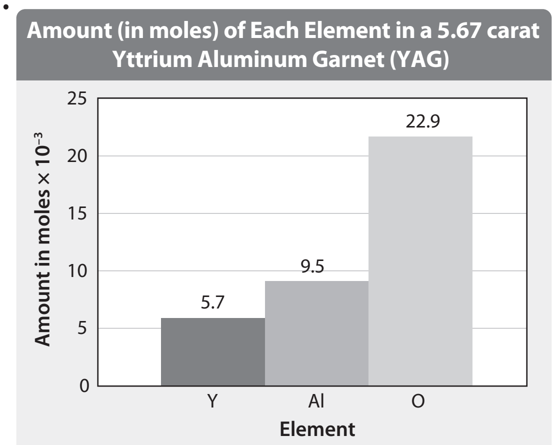


29. Sample answer:

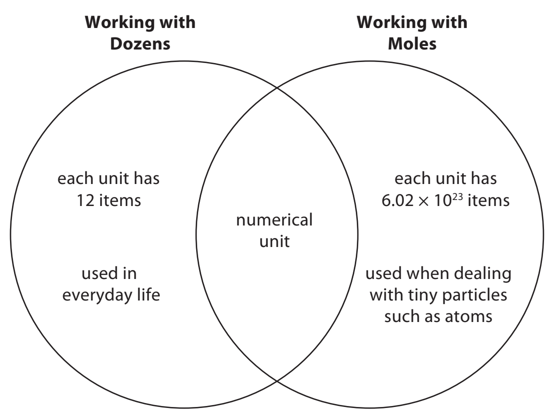


30. There are 21 atoms in each *molecule* of TNT, not in each mole. A mole is a numerical unit referring to 6.02×10^{23} particles. Each mole of TNT contains 6.02×10^{23} molecules, or $21 \times 6.02 \times 10^{23}$ atoms.

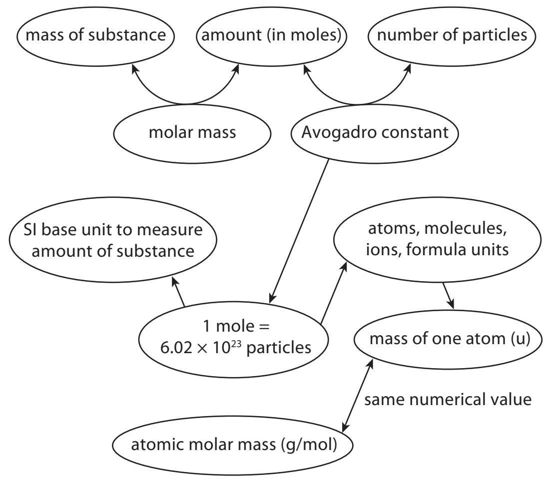
31.



32. Sample answer:



33. Sample answer:



34. a. If the unit mass is available, doctors and pharmacists can perform the necessary calculations to determine correct dosage.

b. Medication doses are specific to the performance of the substance in the body, and may depend on the person's age and body mass. If too much medication is taken, it could lead to serious consequences, including death.

c. If children take ASA during a viral illness, they risk of contracting a rare but serious condition called Reye's Syndrome. Ibuprofen does not carry this risk.

35. Ti_2S_3, titanium(III) sulphide

36. a. 1.2×10^{-5} mol

b. Depending on the mass of the tomato, it may contain about 3 g of lycopene.

37. a. 2×10^{-6} mol

b. There are different levels of toxicity. Acute toxicity generally occurs at doses of 25 000 IU per kg of body weight. Chronic toxicity occurs with daily doses of 4000 IU per kg of body weight over 6 to 15 months. The effects can include headache, nausea, vomiting, dry skin, diarrhea, and bone pain.

38. a. 0.42 g

b. 1.8×10^{-3} mol

c. Compounds containing lead increase the risk of lead absorption and lead poisoning, which can lead to retardation of growth, anemia, and possibly death in extreme cases. Health Canada has set a limit of 10 ppm of lead in products that are applied to the face.

39. a. 13 g of potassium nitrate, 0.28 g sodium fluoride

b. 0.12 mol potassium nitrate, 6.5×10^{-3} mol sodium fluoride

c. Fluoride helps strengthen teeth, and potassium nitrate helps reduce sensitivity of teeth and gums to hot and cold. However, fluoride can be toxic if swallowed. For this reason, young children should be supervised when brushing teeth. In addition, potassium nitrate is known to cause problems in aquatic environments.

40. In the Launch Lab, heat was released in one corner of the bag and gas was produced. On a large scale, if this reaction occurred in a confined space, the release of gas could cause an explosion. Also, the release of large amounts of heat could cause equipment meltdown or a fire. So, industries must be very careful that reactions that release heat or gas are carefully controlled.

Chapter 5 Self-Assessment Questions

(Student textbook pages 254–5)

1. e

2. d

3. b

4. c

5. c

6. d

7. a

8. c

9. d

10. b

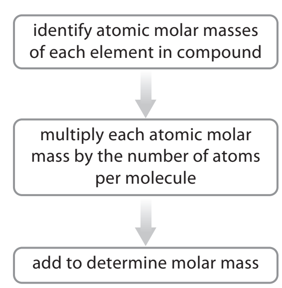
11. Both the atomic molar mass and the molecular molar mass involve one mole of particles. However, the atomic molar mass involves one mole of atoms, while the molecular molar mass involves one mole of molecules.

12. 0.52 g

13. 5.27×10^{-4} mol

14. 3.80×10^{2} g

15. a. Sample answer:



b. 192.1 g/mol

16. 1.09×10^{21} molecules

17. 1.69×10^{22} carbon atoms; 9.87×10^{21} hydrogen atoms; 4.23×10^{21} chlorine atoms; 2.82×10^{21} oxygen atoms

18. a. 111.11 g/mol

 b. 1.20×10^{25} carbon atoms; 1.20×10^{25} hydrogen atoms; 4.82×10^{24} oxygen atoms; 2.41×10^{24} nitrogen atoms

19. The order by mass is: carbon, oxygen, nitrogen, hydrogen. The order by number of atoms is: hydrogen, carbon, oxygen, nitrogen. The orders are not the same, because atoms of different elements have different masses.

20. To convert number of particles of a substance to amount of moles, divide by the Avogadro constant. To convert amount of moles of a substance to number of particles, multiply by the Avogadro constant.

21. a. 4.25 kg b. 14.5 kg

22. They would all have the same number of molecules, because each substance has the same number of molecules per mole (the Avogadro constant).

23. 57.51 g ethanol, 22.49 g water

24. a. Measure the mass of the one dozen marbles and divide by 12. Note: Students may, for simplicity, assume the mass of the plastic bag is negligible, or add a step to weigh a sample of plastic bags and determine an average mass to be subtracted from the mass of the sample.

 b. First method—Multiply the mass of one marble by 365.
 Second method—Divide 365 by 12 to determine how many dozen marbles are in the jar. Multiply the number by the mass of one dozen marbles.

 c. First method—Multiply the mass of one marble by 6.02×10^{23}.
 Second method—Divide 6.02×10^{23} by 12 to find how many dozen marbles are in one mole. Then, multiply the number of dozen marbles by the mass of one dozen to get the mass of one mole of marbles.

 Measuring the mass of one dozen marbles is like calculating the molecular molar mass of a substance with 12 atoms per molecule. To get the mass of one molecule, you would divide the molar mass by the Avogadro constant, just like you divide the mass of the dozen marbles by 12 to get the mass of one marble.

 To find the specific mass of a particular number of particles of a substance, you would divide the number of particles by the Avogadro constant to get the amount of moles, and then multiply by the molar mass to determine the mass.

25. 119 pg/day

Chapter 6 Proportions in Chemical Compounds

Learning Check Questions

(Student textbook page 262)

1. Yes, water is the same regardless of the source. We know this because of the law of definite proportions.

2. Elements will form bonds with other elements until all of their valence shells are full. This can be achieved by many proportions of the same elements.

3. Both sets of terms (definite and constant, and proportion and composition) indicate that the ratio of elements in a compound remains the same.

4. No, the mass percent of carbon in carbon dioxide cannot change, based on the law of definite composition.

5. Yes, because the carbon is present in different ratios in each compound.

6. a. 57.1%

 b. Carbon and oxygen have different mass percents because their molar masses are different.

(Student textbook page 270)

7. There is one atom of carbon for every four atoms of hydrogen, so the ratio is 1:4.

8. Both provide information about the proportions of elements found in the compound. However, the proportions in a percentage composition are based on the overall mass of each element found in the compound, while the proportions in a molecular formula are based on the number of atoms of each element in the compound.

9. The molecular formula shows the actual number of atoms of each element in a molecule of the substance, while the empirical formula only tells you the ratios of the atoms in a molecule.

10. The molecular and empirical formulas are the same when the actual number of atoms of each element in a compound is already in the lowest whole number ratio.

11. Every compound has its own set of unique properties that is a direct result of its structure and composition. A molecular formula reflects the specific composition of a compound, while an empirical formula can represent two or more compounds with the same lowest whole number ratio. For example, NO_2 and N_2O_4 have the same empirical formula (NO_2), but different properties.

12. NO_2 and N_2O_4, because N_2O_4 is a whole-number multiple of NO_2.

(Student textbook page 277)

13. A molecular formula includes the actual numbers of each atom in one molecule of substance. In an empirical formula, although the numbers of each atom are in the correct ratios, they may not be the actual numbers of atoms in one molecule of substance.

14. Choose a standard mass for the substance, such as 100 g. Next, use the given mass percents to calculate the mass of each atom in the 100 g sample. Then use the molar mass of each atom to determine the amount of moles of each atom. Finally, divide or multiply each amount of moles by the correct factor to convert each number to the smallest whole number possible.

15. The molar mass is determined experimentally, usually using a mass spectrometer.

16. A hydrate has a number of H_2O units attached to each molecule, while the anhydrous form does not.

17. Water molecules add mass to a solid. This extra mass can affect measurements and calculations.

18. Heat the substance to drive off all water molecules in the hydrate. Once all the water has been driven off, calculate the difference between the initial and final mass.

Caption Questions

Figure 6.2 (Student textbook page 258): H_2O, or two hydrogen atoms and one oxygen atom

Figure 6.4 (Student textbook page 262): The percentage composition provides the basic information for determining the molecular formula for citric acid. Starting from the formula, scientists can create a plan for synthesizing the compound.

Figure 6.10 (Student textbook page 276): The numbers of atoms in the molecular formula are x times as great as the corresponding numbers in the empirical formula.

Section 6.1 Review Questions

(Student textbook page 267)

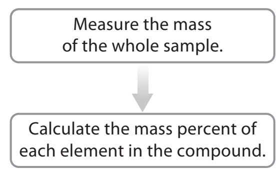
1. They are the same. According to the law of definite proportions, a compound has the same chemical formula regardless of the source of the compound.

2. One mole of a compound is equivalent to the molar mass of the compound. Therefore, the molar masses of elements can be determined directly from the periodic table.

3. Carbon and hydrogen have different molar masses. Therefore, the mass percent of each of these elements is different, even though the number of atoms of each element in a molecule is the same.

4. 42.10% C, 6.49% H, 51.41% O

5. Sample answer:



6. 36% Ca, 64% Cl

7. 0.32% H, 57.95% Au, 41.73% Cl

8. The mass percent of an element in a compound is the ratio of the mass contribution of the element to the total mass of the compound, multiplied by 100. The percentage composition of a compound is a list of the mass percent of each element that makes up the compound.

9. 3.092% H, 31.602% P, 65.306% O

10. 27.74% Mg, 23.57% P, 48.69% O

11. 7.52 g

12. **a.** 40.04% Ca, 12.00% C, 47.96% O
    b. The formation of carbon dioxide gas removes carbon and oxygen from the statue. Therefore, the mass percent of carbon and oxygen in the statue decreases.

13. C

14. Na^+

Section 6.2 Review Questions

(Student textbook page 279)

1. Because only the relative amounts (the simplest ratio) of elements in a compound are required, and not the actual amounts, you can use any mass in the early calculations of the empirical formula.

2. Empirical formula calculations involve a percentage composition. Each percent means "a number out of 100." Using a mass of 100 g reduces the number of conversions needed in a calculation.

3. SnO_2

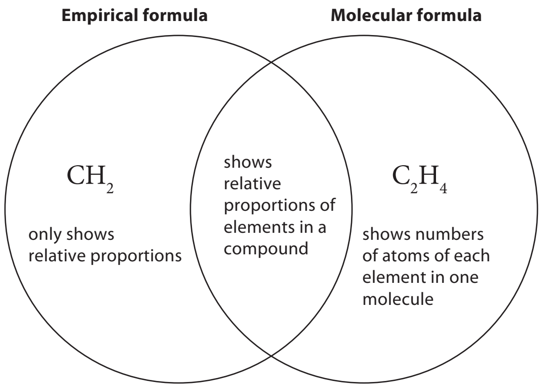
4. $AlCl_3$

5. $KMnO_4$

6. As_2O_3; Percentage Composition of As_2O_3

7. $PbCl_2$

8. Sample answer:



9. 89.16%

10. 7

11. 6

12. You could argue that the evidence is inadmissible, based on the fact that it is an empirical formula, which could belong to any number of different substances, and not the specific molecular formula for the banned substance in question.

13. a. CF_2 **b.** twice as much

14. $C_{10}H_{16}N_2O_8$

15. C_3H_6O

16. $C_{24}H_{24}O_6$

Quirks & Quarks

(Student textbook page 280)

1. Methane hydrate is a frozen substance that is a combination of natural gas and water. Although it looks like ice, methane hydrate burns if you light it with a match.
Benefits—If clean recovery techniques can be found to access the methane hydrate energy reserves, this could supply the world with a much-needed energy source that is cleaner than oil.
Risks—Methane gas is about 24 times as potent a greenhouse gas as carbon dioxide. If the methane that is locked up in these gas hydrates is released into the atmosphere, it could have significant effects on global warming and climate change.

2. With developments in technology, the use of methane hydrates for fuel will likely become feasible. However, at this stage, there are many obstacles to using this fuel, including environmental risk, technical difficulties, and low marketability. Canada is one of the main countries researching the commercial use of methane hydrates. A Japanese research project also exists to develop methane hydrates commercially by 2016.

3. Sample answer: A marine geologist could research locations containing (or likely to contain) methane hydrates. Possible technologies used to harvest methane hydrates include drilling equipment and water pumps.

STSE Chemistry Connection

(Student textbook page 281)

Sample answer: In a 2006 study called "Differentiating Writing Inks Using Direct Analysis in Real Time Mass Spectrometry," scientists used a mass spectrometer to examine a variety of inks. The results were placed in a database. When tested, samples could be matched with the exact type of ink using this database. This information is useful as a way to establish the authenticity and/or age of a document.

Practice Problems

(Student textbook page 260)

1. 22.27%

2. 69.55%

3. 7.8%

4. 53.28%

5. 25.6%

6. 27%

7. dichromic acid

8. sulfurous acid

9. 63.89%

10. $ZnS(s)$, $Cu_2S(s)$, $PbS(s)$

(Student textbook page 264)

11. 82% N; 18% H

12. 68.4% Cr; 31.6% O

13. 40.0% C; 6.7% H; 53.3% O

14. 48% Ni; 17% P; 35% O

15. 37.0% C; 2.20% H; 18.5% N; 42.3% O

16. 67.10% Zn; 32.90% S

17. 127.8 g Cu; 32.2 g S

18. 24.74% K; 34.76% Mn; 40.50% O

19. 10.11% C; 0.80% H; 89.09% Cl

20. No, the percentage composition of carbon in the sample is 64.8%. If the sample were ethanol, the percentage composition of carbon would be 52.1%.

(Student textbook page 266)

21. 63.14% Mn, 36.86% S

22. 93.10% Ag, 6.90% O

23. 2.06% H, 32.69% S, 65.25% O

24. 34.59% Al, 61.53% O, 3.88% H

25. 41.40% Sr, 13.24% N, 45.36% O

26. 73.27% C, 3.85% H, 10.68% N, 12.20% O

27. 205 kg

28. 127 kg

29. 17.1 g

30. 248 kg

(Student textbook page 273)

31. CH_3

32. Mg_2Cl

33. $CuSO_4$

34. $K_2Cr_2O_7$

35. NH_3

36. Li_2O

37. BF_3

38. Cl_3S_5

39. Na_2CO_3

40. P_2O_5

(Student textbook pages 275–6)

41. $C_6H_{12}O_6(s)$

42. $C_8H_{10}(\ell)$

43. $C_4O_2H_{10}(\ell)$

44. $C_8H_8(s)$

45. $HgCl(s)$

46. $C_8H_{10}N_4O_2(s)$

47. $C_2H_5NO_2(s)$

48. $B_2H_6(g)$

49. a. C_3ONH_8 **b.** $C_6O_2N_2H_{16}$

50. Its empirical formula is $C_9H_{12}O$ and its molecular formula is $C_{18}H_{24}O_2$.

(Student textbook page 278)

51. 50.88%

52. 62.97%

53. 20.93%

54. 62.97%

55. $MgSO_4 \bullet 7H_2O(s)$, $Ba(OH)_2 \bullet 8H_2O(s)$
$CaCl_2 \bullet 2H_2O(s)$, $Mn(SO_4)_2 \bullet 2H_2O(s)$

56. 5

57. 4

58. $Cr(NO_3)_3 \bullet 9H_2O$

59. $MgI_2 \bullet 8H_2O$

60. 2.83 g

Chapter 6 Review Questions

(Student textbook pages 289–91)

1. b

2. d

3. a

4. d

5. c

6. e

7. a

8. b

9. b

10. a. 1:1 **b.** 1:2:1 **c.** 3:1:4 **d.** 1:1:3

11. It does not provide more information when the actual amounts of each element inside a compound are already the lowest whole number ratio.

12. Row 1—empirical formula; percentage composition; none; yes; yes; no
Row 2—percentage composition; pure substance; which elements are present, the mass of the sample, and the mass of each element in the sample; yes; no, only relative masses; no
Row 3—molecular formula; empirical formula; molar mass; yes; yes; yes

13. An empirical formula is the relative ratio of elements in a compound, but not the actual amount of moles or atoms. For this reason, it is the simplest ratio of elements in a compound.

14. The percentage composition or the empirical formula, and the molar mass.

15. Yes, according to the law of definite proportions.

16. No, you need the percentage composition or empirical formula.

17. Yes, a molecular formula indicates the number of atoms in one molecule (or formula unit) of a compound.

18. Because the empirical formula gives the basic ratio of elements in a compound, it could identify several compounds that have the same ratio of elements but different numbers of atoms in a molecule—giving them very different properties.

19. $MgCl_2$

20. $HBrO_3$

21. $C_6H_{12}O_6$

22. $C_4H_6O_6$

23. $C_{20}H_{40}O_2$

24. $C_{12}H_{12}Cl_9F_6$

25. $C_6H_6O_2$

26. CH_2

27. 43.09%

28. 80.48 g/mol

29. barium chloride dihydrate, $BaCl_2{\bullet}2H_2O$

30. hematite

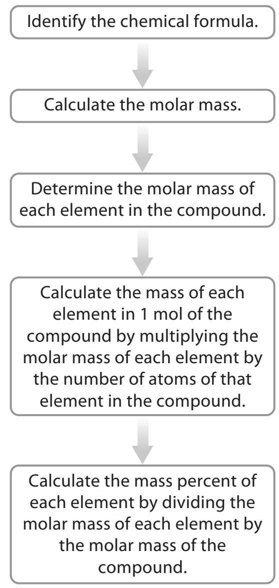
31. 49.47% C, 5.20% H, 28.85% N, 16.48% O

32. **a.** $Ba(OH)_2{\bullet}8H_2O$; 45.6%
    b. $Na_2CO_3{\bullet}10H_2O$; 62.9%
    c. $CoCl_2{\bullet}6H_2O$; 45.4%
    d. $FePO_4{\bullet}4H_2O$; 32.3%
    e. $CaCl_2{\bullet}2H_2O$; 24.5%

33. $C_4H_4O_2$

34. No, the molecular formula must be a whole number multiple of the empirical formula.

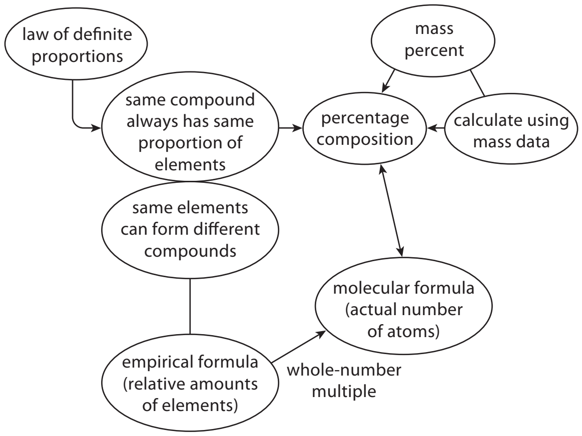
35. Sample answer: For starting from mass data see answer to question 5 of Section 6.1 Review. Starting from a chemical formula:



36. Presentations could explain that when determining percentage composition values, one mole of a compound is the starting sample amount, and the mass of the elements are whole-number multiples of their molar mass values.

37. Sample answer: Measure the mass of the hydrated sample before your procedure, so that you know the mass of the compound including water molecules. Next, heat the hydrated sample until the sample is in anhydrous form (e.g., crystals change colour). Find the mass of the anhydrous sample. This is the mass of the compound without water. Determine the amount of moles of the anhydrous compound and the molar mass of the compound. Determine the amount of moles of water using the mass of the water (the difference between the first and second sample masses) and the molar mass of water. Now determine the relative amount of moles of water per mole of the compound. This is the same as the number of molecules of water per formula unit of the compound. Use this information to write the formula for the hydrate.

38. Sample answer:



39. Fe_3O_4

40. a. Trial 1: 80.4% Zn, 19.6% O

Trial 2: 80.4% Zn, 19.6% O

Trial 3: 86.2% Zn, 13.8% O

Trial 4: 80.4% Zn, 19.6% O

b. Discard Trial 3. Three of the trials give whole number ratios, but Trial 3 seems to have a lower relative final mass and does not give whole number ratios.

c. The reaction was incomplete. Not all of the Zn reacted and the sample did not absorb as much oxygen. As a result, the overall mass of the final sample is too low.

d. $x = 1, y = 1$

41. a. 64.67% NiS, 43.93% NiAs, 30.47% $(Ni,Fe)_9S_8$; Sample answer: I would mine the millerite formation first, because it has the highest percent by mass of nickel.

b. Emissions of sulfur dioxide from mining operations created acid rain conditions that devastated the regions' ecology. Since then, the region has gone through several changes to mining and emission policies, and the ecology has recovered greatly as a result.

42. 11.50 g, 12.99 g, 15.99 g, 17.48 g, 18.98 g, 20.48 g, 26.46 g

Chapter 6 Self-Assessment Questions

(Student textbook pages 292–3)

1. a

2. c

3. e

4. d

5. d

6. b

7. e

8. b

9. d

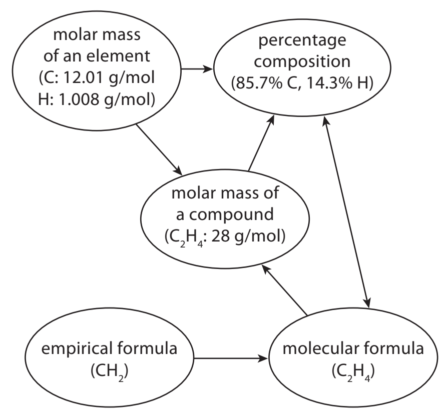
10. d

11. According to the law of definite proportions, pure samples of a compound always have the same percentage composition. For example, water is always H_2O, regardless of the source or state.

12. Sample answer:

Empirical Formula	Molecular Formula
• simplest formula • shows the lowest whole number ratio of elements in the compound • e.g., empirical formula for benzene is CH	• actual formula for a compound • shows number of atoms of each element that make up a molecule • e.g., molecular formula for benzene is C_6H_6

13. Sample answer:



14. $BaCO_3$, Percentage Composition of $BaCO_3$

15. H_2SO_3

16. N_2O_3

17. $NaNO_2$

18. 7

19. C_8H_{20}

20. N_2O_2

21. 68.9%

22. a. $SiCl_3$ **b.** Si_2Cl_6

23. a. NO_2 **b.** N_2O_4

24. a. Sodium carbonate heptahydrate

b. 54.34% **c.** 52 kg

25. Chalcocite, Cu_2S, because this ore contains a higher percent by mass of copper.

Chapter 7 Chemical Reactions and Stoichiometry

Learning Check Questions

(Student textbook page 299)

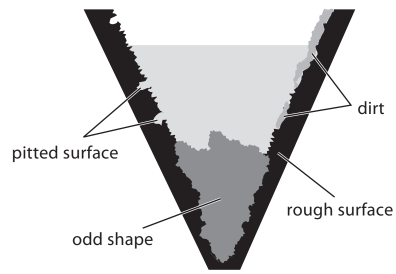
1. 10 slices of toast, 10 turkey slices, 5 lettuce leaves, and 5 tomato slices
2. The exact proportion of moles of each reactant and product is needed to determine the relative amounts of reactants and products in a complete reaction.
3. The coefficients show the reacting molecular and mole ratios. One mole of methane reacts completely with two moles of oxygen to produce one mole of carbon dioxide and two moles of water.
4. You can determine the relative amount of moles of each reactant and product in a complete reaction.
5. Because we are interested in the mole ratio for each reactant and product, not for each individual atom. Coefficients refer to relative amounts of entire molecules, while subscripts refer to the relative amounts of each atom within a molecule.
6. **a.** 2 mol C_2H_6(g) : 4 mol CO_2(g), or 1:2
   b. 2 mol C_2H_6(g) : 7 mol O_2(g), or 2:7
   c. 4 mol CO_2(g) : 6 mol H_2O(g), or 2:3

(Student textbook page 309)

7. The exact molar amount of a reactant or product, as predicted by a balanced chemical equation.
8. **a.** The limiting reactant is gas, and the excess reactant is oxygen in the air. I assumed the reaction took place in normal air
   b. The limiting reactant is deposits ($CaCO_3$(s)), and the excess reactant is vinegar. I assumed lots of vinegar was used.
   c. The limiting reactant is potato, and the excess reactant is oxygen in the air. I assumed the reaction took place in normal air.
9. tomato
10. Not necessarily. The limiting reactant is the reactant with an insufficient stoichiometric amount based on the balanced chemical reaction. In other words, it is the reactant that would be used up while the other reactants are still available. For example, if the mole ratio is high (e.g., 9 mol A : 1 mol B), then reactant A is limiting even if 8 mol of A are available for each mol of B.
11. The amount in excess is not used in the reaction.
12. Oxygen, because I assume that there will be plenty of oxygen remaining in the air after all the phosphorus has reacted.

(Student textbook page 316)

13. The theoretical yield is the mass or amount of product calculated using the chemical equation, the mole ratios, and the mass of reactants. The actual yield is the mass or amount of product obtained experimentally.
14. The theoretical yield is usually higher, since some product is usually lost during the experiment.
15. Improper lab techniques may reduce reaction yields in a number of ways. Some product may cling to the lab equipment and not be properly rinsed and collected. Spillage might occur. Measurements might not be made correctly.
16. Sample answer:



17. If a reactant is not pure then the actual mass of reactant is less than the measured mass.
18. Sample answer: A leak in a tank will cause the amount of fluid in the tank to decrease. A competing reaction drains away some of the desired product just like the leak in the tank.

Caption Questions

Figure 7.1 (Student textbook page 296): 8 slices of toast, 8 turkey slices, 4 lettuce leaves, and 4 tomato slices

Figure 7.3 (Student textbook page 297): 20 atoms of H, 10 atoms of O

Section 7.1 Review Questions

(Student textbook page 305)

1. relative molar amounts
2. Stoichiometric calculations require mole ratios for reactants and products. These ratios are obtained from a balanced chemical equation.
3. **a.** 4:3:2, 4:3, 4:2, 3:2
   b. 3:4:1:4, 3:4, 3:1, 4:1, 4:4
   c. 2:2:1, 2:2, 2:1

d. 2:1:2, 2:1, 2:2

e. 1:1:1, 1:1

4. 6.0×10^2 g

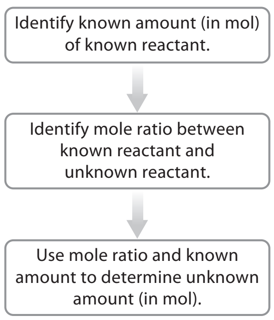
5. 28 g

6. The student is assuming that 1 mol of each element/compound has a mass of 1 g, which is not true. The correct masses are Mg(s): 24.305 g, and Cl_2(g): 70.906 g.

7. a. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

b. 526.2 kg or about half a tonne

8. Sample answer:



9.

	$2HBr(aq)$	$Ca(OH)_2(aq)$	$CaBr_2(aq)$	$2H_2O(\ell)$
Amount (mol)	2	1	1	2
Number of Units	1.20×10^{24}	6.02×10^{23}	6.02×10^{23}	1.20×10^{24}
Mass (g)	162 g	74.0 g	200 g	36.0 g

10. a. 1.7×10^{22} formula units

b. 2.30×10^{22} formula units

11. Sample answer: I have noticed that excess fertilizer from your lawn is being washed into Lake X. Since algae growth is directly linked to fertilizer contamination, your excess fertilizer is causing an exponential growth of algae in the lake. This growth is using up dissolved oxygen in the lake, and is harming aquatic organisms as a result. In addition, your lawn does not benefit from excess amounts of fertilizer, and it is costing you extra money. Please consider calculating the correct amount of fertilizer needed by your lawn and using only that amount. You will save money, your lawn will be adequately maintained, and you will protect the ecosystem in Lake X.

12. Sample answer: Notice to Gardeners in the X Area: Please be careful to use the smallest quantities of fertilizer, pesticides, and herbicides that are appropriate for the size of your garden. Using excess amounts will not benefit your garden, and harms local ecosystems.

Section 7.2 Review Questions

(Student textbook page 313)

1. water

2. O_2(g)

3. O_2(g)

4. a. $FeCl_3$(aq)

b. 2.83 g NaOH(aq) remains

c. 37.8 g of $Fe(OH)_3$(s) and 62.0 g of NaCl(s) form

5. 3.47 g

6. A limiting reactant is a reactant that is present in insufficient amount as compared to the other reactants of a chemical reaction. For example, a bonfire will never use up all the oxygen in my backyard, because much less wood is present than oxygen.

7. Sample answer:
(1) Write and balance the chemical equation.
(2) Identify the mole ratio for the reactants.
(3) Determine the quantity in moles of each reactant.
(4) Use the mole ratio to compare the amount of moles of each reactant and the amount of product that each would produce.
(5) Identify the limiting reactant.

8. Based on stoichiometry, there is a direct relationship between the amount of oxygen and the amount of wax that can react in a combustion reaction. When the cover is on the container, the oxygen in the container becomes the limiting reactant. When all the oxygen is used up, the flame stops burning. However, if the lid is removed, there is plenty more oxygen available, and the oxygen is no longer the limiting reactant. Now the flame will burn until the wax (the limiting reactant) is gone.

9. 15.75 g

10. a. The graph levels off because at that point, the sodium chloride is all used up, and no more silver chloride is produced, no matter how much silver nitrate is added.

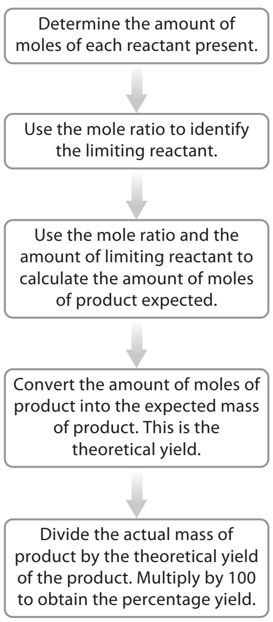
b. 0.15 mol

11. At first, I did not add enough vinegar to react with all the deposits. The vinegar was the limiting reactant. Then, when I added more vinegar, the vinegar became the excess reactant and the deposits became the limiting reactant. All the deposits reacted with the vinegar.

Section 7.3 Review Questions

(Student textbook page 321)

1. As long as you use the same units for the theoretical yield and the actual yield, it does not matter which units you use. A percent does not have units.
2. The theoretical yield is the amount of product you expect, based on the amount of reactant. The actual yield is the amount of product you actually get. For example, if I make a cookie recipe, I might expect 12 cookies. I might only get 10 cookies, though, if I make bigger cookies or do not scrape the bowl.
3. Sample answer: If a reaction rate is very slow, the reaction may not be complete by the time I measure the actual yield.
4. 94.60%
5. 91.9%
6. You would assume that the reactants are 100% pure.
7. Sample answer:



8. Sample answer: If vinegar is used instead of pure acetic acid, the reaction solution will be very dilute. As a result, the reaction may proceed more slowly. Also, you will need to use a much larger quantity of vinegar.
9. If a company is producing a chemical compound for product, the company wants to ensure that the reactants are used up and not wasted. A reaction with a low percentage yield means that reactants are being wasted, and the company is losing money. A high percentage yield means that the reaction is efficient and money is not being wasted.
10. Sample answer:
    a. use a non-absorbent stir stick
    b. do not transfer reactants between vessels
    c. breeze may blow some dried reactant off the evaporating dish
    d. rinse the graduated cylinder and add the wash to the mixture
11. **a.** 3.4 g of I_2, 1.6 g of NaCl
    b. 1.0 g of NaCl
    c. Either a yes or no answer is acceptable. Sample answers: No, it is not reasonable because the scientist collecting the products might spill one product on the floor.
    Yes, it is reasonable since any decrease in yield that affects the efficiency of the reaction as a whole should likely have the same effect on both products.
12. Theoretical yield: 2.638 g $CuSO_4(s)$, 1.4890 g H_2O
    Actual yield: 2.913 g $CuSO_4(s)$, 1.214 g H_2O
    Percentage yield: 110.4% $CuSO_4(s)$, 81.5% H_2O

Practice Problems

(Student textbook page 298)

1. 2 mol Mg(s) : 1 mol $O_2(g)$: 2 mol MgO(s)
2. 2 mol NO(g) : 1 mol $O_2(g)$: 2 mol $NO_2(g)$
3. 1 mol Ca(s) : 2 mol $H_2O(\ell)$: 1 mol $Ca(OH)_2(s)$: 1 mol $H_2(g)$
4. 2 mol $C_2H_6(g)$: 7 mol $O_2(g)$: 4 mol $CO_2(g)$: 6 mol $H_2O(g)$
5. 5 molecules
6. 155 molecules of $AlCl_3(s)$
7. 3.4×10^{23}
8. 3.9×10^{24}
9. 2.72×10^{24}
10. 2

(Student textbook pages 300–1)

11. 0.25 mol
12. 6.00 mol
13. 4.50×10^4 mol
14. 3.6 mol
15. 4.70 mol
16. **a.** 46.8 mol **b.** 187 mol
17. 56.5 mol
18. 6.45 mol
19. 5.1×10^{23} molecules
20. 7.24×10^5 mol

(Student textbook page 304)

21. 1.3×10^5 g

22. 860 kg

23. 726 g

24. 0.123 mol

25. 0.421 g

26. 4.35 g

27. 10.3 g

28. 0.963 g

29. 5.16×10^{-2} g

30. $Al_4C_3(s) + 12H_2O(\ell) \rightarrow 4Al(OH)_3(s) + 3CH_4(g)$; 21.0 g

(Student textbook page 309)

31. $CaF_2(s)$

32. $C_7H_6O_3(aq)$

33. $H_2O(\ell)$

34. $NiCl_2(aq)$

35. $HNO_3(aq)$

36. Li(s) is limiting, $O_2(g)$ is in excess

37. $Na_2S_2O_3(aq)$

38. $C_3H_6(g)$

39. 61 g

(Student textbook page 311)

40. a. $O_2(g)$ **b.** 8 mol

41. a. $AgNO_3(aq)$ **b.** 0.03519 mol

42. 66.98 g

43. 11.4 g

44. 172 g

45. 9.8 g

46. 57.3 g

47. 11.8 g

48. 69.94 g

49. a. 0.446 g **b.** $F_2(g)$, 24.0 g

50. a. The total mass of $CaCl_2(aq)$ is 776.9 g. (The excess mass is 388.4 g.)

b. The mass of $AgNO_3(aq)$ is 1189 g.

c. 1003 g of AgCl(s), and 574.3 g of $Ca(NO_3)_2(aq)$

(Student textbook page 319)

51. 0.97 g, 0.77 g

52. a. 37.3 g **b.** 75.6%

53. 79.3%

54. a. 8.32 g **b.** 5.99 g

55. 44.6 g

56. 11.9 g

57. 61.8 g

58. 0.46 g

59. a. 1.51 kg **b.** 1180 g

60. 35 g

Chapter 7 Review Questions

(Student textbook pages 331–3)

1. b

2. d

3. c

4. b

5. d

6. a

7. a

8. e

9. Sample answer: Both a cooking recipe and a balanced chemical equation give you information on the substances and the amounts needed to react together to produce a specific product. However, a balanced chemical equation gives much more information than a cooking recipe, such as information about the number of particles involved, and specific information about all the products.

10. a. vinegar is the limiting reactant, and baking soda is the excess reactant

b. vinegar is the limiting reactant, and the deposits are the excess reactant

c. the tablet is the limiting reactant, and water is the excess reactant

11. The coefficients tell you how many moles will react or form, relative to the amount of moles of other reactants or products.

12. The percentage yield will be lower, because the actual amount of limiting reactant is smaller than what is measured.

13. By applying stoichiometry, manufacturers can save money. For example, by making sure that the most expensive reactant is the limiting reactant, manufacturers can ensure that none of that reactant is wasted.

14. First, I would write a balanced chemical equation and determine the appropriate mole ratio. Then I would use the mole ratio and the amount in moles of the first reactant to calculate the amount in moles of the other reactant.

15. First, I would write a balanced chemical equation and determine the appropriate mole ratio. Then I would use the mole ratio and the number of molecules of reactant to calculate the number of molecules of product. (Because the mole ratio is a ratio, it applies to molecules as well as to amount of moles.)

16. Answers should specify that they are referring to amount of moles and not to mass. Sample answer: This reaction requires three times as many moles of hydrogen as nitrogen to proceed.

17. a. $2Al(s) + 3Br_2(g) \rightarrow 2AlBr_3(s)$

b. 7.5 mol **c.** 5.0 mol

18. Theoretical yield: 14.77 g, percentage yield: 98.75%

19. 2.3 mol

20. 0.488 g, 93.2%

21. It could be a dilute solution, in which case, the $AgNO_3$ is limiting and "completion" occurs when all the $AgNO_3$ is used up. Or, it could be a concentrated solution, in which case the copper wire is limiting and "completion" occurs when all the copper is used up.

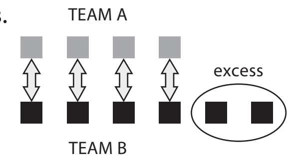
22. a. No, a comparison of the amount of moles of lithium nitride and of water reveals that there is not three times more water than lithium nitride present. The water is the limiting reactant, so the lithium nitride will not react completely.

b. 3.86 g

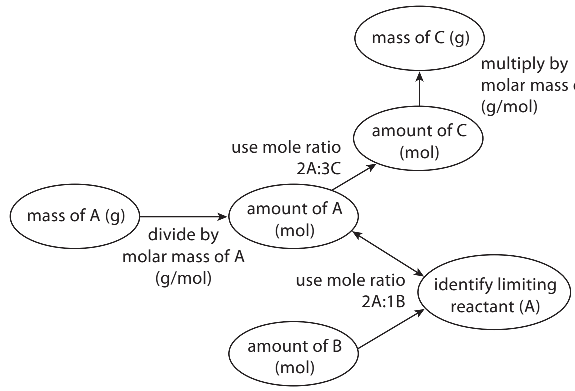
23. Sample answer: The mass of oxygen yielded could be inferred from the balanced chemical equation following the steps below, or the oxygen could be collected and measured.
(1) Use the mass of potassium chlorate and the balanced chemical equation to calculate the theoretical yields of potassium chloride and oxygen gas.
(2) Carry out the reaction.
(3) Determine the mass of the potassium chloride product. Subtract this mass from the mass of the reactant to obtain the mass of oxygen gas produced (law of conservation of mass).
(4) Use the mass of the potassium chloride to calculate the actual yield. Then calculate the percentage yield of potassium chloride.
(5) Use the mass of oxygen gas to calculate the actual yield. Then calculate the percentage yield of oxygen gas.

24. 87.3%

25. Sample answer: If two school wrestling teams are competing, the individuals pair up for each match. However, if Team A has fewer people than Team B, there can only be as many pairs as the number of people in Team A. Team A is the limiting reactant, and Team B is in excess.

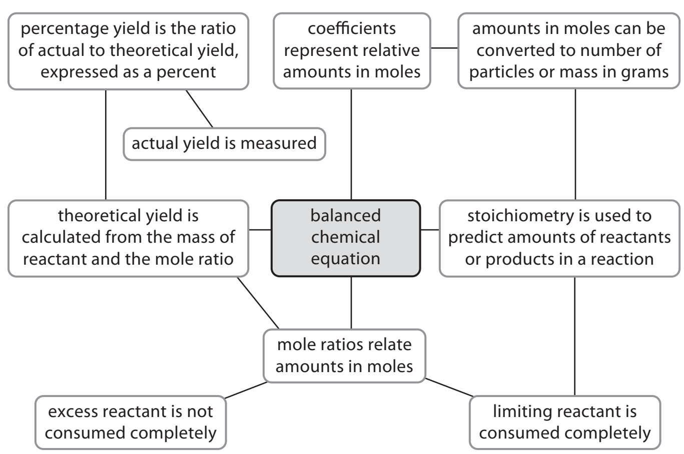


26. Sample answer:



27. Sample answer: At Company X, our planning protects the environment while benefitting our clients. By carefully calculating the appropriate quantities of chemicals to use in each reaction batch, we reduce our chemical waste to prevent harm to the environment. This also saves money, so we can keep our prices low.

28. Sample answer:



29. Sample answer: When washing my clothes, I measure out appropriate amounts of detergent and bleach for the volume of water in the washing machine.

30. Sample answer: In Canadian gold mining, gold ore may be processed using cyanide to extract the gold. Cyanide is extremely toxic. Using too much cyanide to extract the gold may result in toxic waste that can leak into the environment.

31. a.

Type of Fuel	Coefficient			
	Fuel	O_2	H_2O	CO_2
hydrogen	2	1	2	–
gasoline	2	25	18	16
propane	1	5	4	3
methane	1	2	2	1

b. Burning hydrogen fuel does not produce carbon dioxide. However, at present, most energy- and cost-efficient ways to produce hydrogen fuel use fossil fuels, which release carbon dioxide. If a reliable, clean, and economical way of producing hydrogen fuel is developed, then using hydrogen fuel will reduce human production of carbon dioxide, and help limit climate change.

32. Calcium chloride is much cheaper than silver nitrate, and less toxic. I choose calcium chloride to be in excess. This will ensure that no silver nitrate is wasted or needs to be disposed of.

33. Nitrogen oxides are produced when burning fossil fuels in air. These include nitric oxide, NO(g), and nitrous oxide, N_2O(g). Nitric oxide is an air pollutant, and nitrous oxide is a greenhouse gas that contributes to climate change.

34. a. The word *tar* commonly implies something dark, sticky, and impure or contaminated. It may have a negative connotation. The word *oil* sounds cleaner and more useful. Referring to *oil sands* rather than *tar sands* may spread the misconception that this is a relatively clean and easy source of energy.

b. Extracting fossil fuels from tar sands requires a large amount of energy. Because of the low percentage yield, a great deal of this energy is wasted. Most of our energy is produced by burning fossil fuels, so wasted energy means that more greenhouse gases are emitted. Greenhouse gases contribute to climate change.

35. Amount in moles can also be used in stoichiometry. In some cases, volume can also be used.

36. Sample answer: As a concerned citizen and as a chemistry student, I would like to point out that industries commonly use large quantities of excess chemicals to save money by ensuring that industrial reactions go to completion. However, these excess chemicals result in waste and can cause environmental harm when they are disposed of. Please consider creating legislation to limit the quantity of excess chemicals that can be used in an industrial reaction.

Chapter 7 Self-Assessment Questions

(Student textbook pages 334–5)

1. b

2. b

3. a

4. e

5. c

6. a

7. e

8. b

9. a

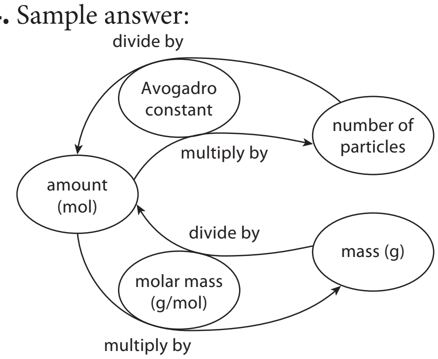
10. The limiting reactant is the reactant that is completely used up in a reaction. The excess reactant is the reactant that has some left over after the reaction is done.

11. Sample answer: If I make cookies out of flour, milk, sugar, eggs, and butter, I may run out of milk after one batch, even though I still have plenty of other ingredients. In this case, the milk is the limiting reactant.

12. 17.81 g

13. a. 4.68×10^{-2} mol **b.** 4.5 mol

14. Sample answer:



Sample calculation: 1 mol of Au contains 6.02×10^{23} particles, has a molar mass of 196.97 g/mol, 196.97 u.

15. 93.2 kg

16. a. $3Zn(s) + 2FeCl_3(aq) \rightarrow 2Fe(s) + 3ZnCl_2(aq)$

b. 272.73 g **c.** 479.12 g

17. Sample answer: (1) Write a balanced chemical equation. Identify the mole ratio you need.
(2) Use the mass and molar mass of the reactant, sodium hydrogen carbonate, to calculate the amount of reactant in moles.
(3) Use the amount of reactant and the mole ratio to calculate the amount of sodium carbonate that will be produced, in moles.
(4) Use the amount of sodium carbonate in moles, with its molar mass, to calculate the theoretical yield in grams.
(5) Carry out the reaction. Measure the mass of sodium carbonate produced.
(6) Divide the mass produced by the theoretical mass, and multiply by 100 to obtain the percentage yield of sodium carbonate.

18. Sample answer: If I make a shirt out of fabric, I may have more fabric than is needed. Once the shirt is done, I will have fabric left over. The fabric is an excess reactant.

19. a. The burning magnesium is emitting smoke, which is probably blowing away. As a result, some of the product will be lost.
b. 55.78%

20. 0.31 g

21. Sample answer: If the student does not rinse off the stirring rod and add the wash to the drying precipitate, some of the solid may stick to the stirring rod and be lost. On the other hand, stirring the precipitate may ensure that no liquid remains to affect the final mass, so it is not a bad idea provided that it is done correctly.

22. 0.185 mol

23. $O_2(g)$

24. The mole ratios are 2:1:2 and 1:1:2. To identify the limiting reactant, compare the amount of moles of each reactant present with the amount of moles of product that would be produced. The reactant that would produce the least amount of product is the limiting reactant.

25. 12.5 mol; $2SO_2(s) + 2H_2O(\ell) + O_2(g) \rightarrow 2H_2SO_4$

Unit 3 Review Questions

(Student textbook pages 339–43)

1. e

2. e

3. c

4. a

5. b

6. e

7. b

8. b

9. Molecular mass refers to the mass of one molecule in a molecular compound, such as carbon dioxide. Formula mass refers to the mass of one formula unit in an ionic compound, such as sodium chloride.

10. You need the percentage composition of the compound, and the molar mass of each element in the compound.

11. An empirical formula tells you the ratio of diferent types of elements present in a compound.

12. a. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
b. 1.07 mol, 2.16 g

13. Both sugars have the same empirical formula, CH_2O. However, they are different compounds. A molecule of one sugar has different numbers of carbon, hydrogen, and oxygen atoms than a molecule of the other sugar. This difference is shown in their molecular formulas

14. 19 kg

15. a. $O_2(g)$ **b.** 19.1 g

16. 17.2 g

17. a. 1.97×10^{23} g **c.** 2×10^{-11} mol
b. 1.66×10^{-22} mol, 32.8 g **d.** 4.13×10^{12} g

18. 27% K, 35% Cr and 38% O

19. 32.37% Na, 22.58% S, 45.05% O

20. a. $Pb(SO_4)_2$, $Pb_2(SO_4)_4$ **b.** lead(IV) sulfate

21. 1.443 g, Sb_2S_3

22. $Na_2C_8H_4O_4$

23. a. $C_3H_5O_2I$ **b.** $C_6H_{10}O_4I_2$

24. a. 40% C, 6.7% H, 53.3% O; CH_2O
b. $C_5H_{10}O_5$

25. a. 36.1%
b. i. If the material is not heated for long enough, some of the water will still remain. The students' calculation of percentage composition will be too low.
ii. If hydrate is spilled on the balance, that mass will be measured but will not take part in the reaction. The calculation will be too high.
iii. If some of the reaction material is removed, then the students' calculation will be too high.

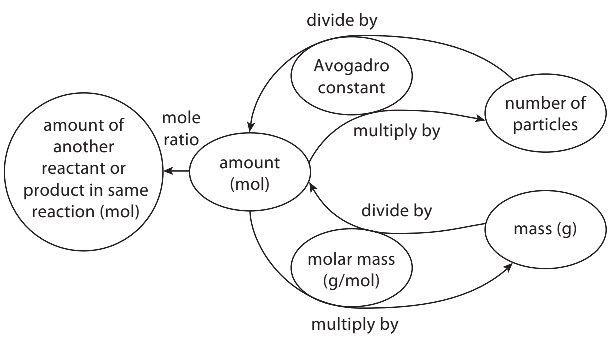
26. Sample answer:
$AgNO_3(aq) + KBr(aq) \rightarrow AgBr(s) + KNO_3(aq)$. In

this reaction, product will likely be lost by remaining on the lab equipment. Other reasons for a lower percentage yield include impure reactants, a slow reaction time, and competing reactions.

27. 0.3 g

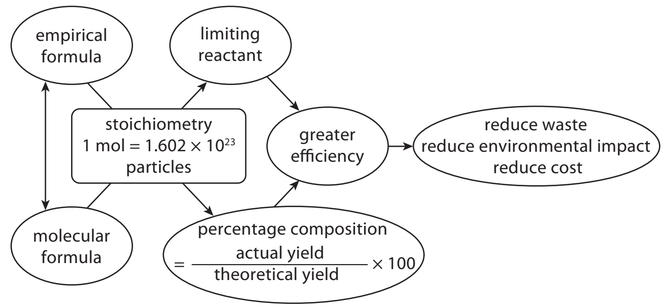
28. Sample answer:
- You have 27.0 g of Al(s) and 79.5 g CuO(s). The reaction produces 53.5 g Cu(s).
- Balance the chemical equation:
  $2Al(s) + 3CuO(aq) \rightarrow Al_2O_3(aq) + 3\,Cu(s)$
- Identify the molar mass of each reactant and of the desired product: 27.0 g/mol, 79.5 g/mol, and 63.5 g/mol
- Calculate the amount of moles of each reactant: 1 mol Al(s) and 1 mol CuO(s)
- Identify the limiting reactant: CuO(s)
- Calculate the amount of Cu(s) that would be produced by 1 mol of CuO(s): 1 mol
- Calculate the theoretical yield of Cu(s) in grams: 63.5 g
- Calculate the percentage yield of Cu(s): 84.25%

29. Avogadro constant, molar mass, and mole ratio; sample answer:



30. Sample answer:
(1) Determine the mass of a sample of the compound.
(2) Place the sample in a heat-proof container of known mass, and heat it on a hot plate for 1 h, or until a physical change occurs.
(3) Determine the mass of the sample after heating. If the mass has not changed, the sample is not a hydrate. If the mass has changed significantly, the sample is probably a hydrate.
(4) If the sample is a hydrate, subtract the new mass from the original mass to calculate the mass of water that was lost.
(5) Use the molar mass of water, 18 g/mol, to calculate the amount of moles of water that were lost.
(6) Use the new mass of the sample and the molar mass of the anhydrous compound to calculate the amount of moles of anhydrous compound present.
(7) Divide the amount of moles of water by the amount of moles of anhydrous compound to identify the ratio of the amount of moles of water per mole of anhydrous compound. This ratio is the same as the number of water molecules per molecule of anhydrous compound.

31. Sample answer:



32. a. $2C_8H_{18}(\ell) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$
b. 0.700 mol; 30.8 g

33. a. Sample answer: Ammonium sulfate is used by farmers around the world to increase crop yield. It leaches into groundwater from the soil, where it may cause eutrophication (excessive algae growth that reduces the amount of oxygen in water systems). However, this compound biodegrades in water, so its effects disappear quickly. I think that the benefits of this fertilizer outweigh the risks.
b. Research may be presented in a wide range of formats including as a video clip, poster, pamphlet, talk, web site, or blog entry.

34. A high percentage yield is important for the chemical industry because it ensures minimal waste, which reduces cost as well as the risk of environmental damage from waste chemicals.

35. HCO_2

36. CuO, copper(II) oxide or cupric oxide

37. C_3H_6O

38. a. 115 kg **b.** 144 kg **c.** 228 kg **d.** 81.6%

39. $Na_2SO_4{\bullet}10H_2O$, sodium sulfate decahydrate

40. 11.99 g Ag per 1.0 g Al

41. a. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
$2C_8H_{18}(\ell) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

b. The fuel to oxygen ratio is much higher for gasoline than for propane (2:25 compared to 1:5). This means that gasoline requires more oxygen than propane in order to undergo complete combustion. In other words, gasoline is more likely than propane to undergo incomplete combustion and produce carbon monoxide.

42. a. 15.1 kg H_2O; two cartridges are needed

b. 51.7 kg $CO_2(g)$

Unit 3 Self-Assessment Questions

(Student textbook pages 344–5)

1. d

2. c

3. b

4. a

5. b

6. d

7. c

8. e

9. b

10. c

11. 11.2 g

12. Sample answer: First, I would help my classmate define the law of definite proportions in his or her own words (e.g., The law of definite proportions means that a chemical compound has the same proportions of elements by mass, no matter where the compound comes from.)

Next, I would provide two or three examples of this law (e.g., A sample of water from your tap contains 2 g of hydrogen for every 18 g of oxygen. So does a sample of water from India, or even from Mars.)

Finally, I would prepare a sample problem illustrating the law for my classmate to complete. I would ask my classmate to explain how this sample problem illustrates the law of definite proportions. (E.g., Identify the proportions of elements by mass in a 1-mol sample of CH_3OH. Use the law of definite proportions to explain why this answer is the same, no matter where the methanol comes from.)

13. CH_2, C_6H_{12}

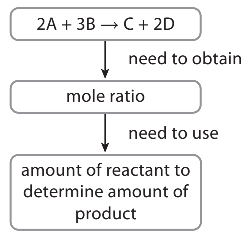
14. A competing reaction might have been occurring. The reaction rate might have been slow, so that the reaction did not go to completion. One or more of the reactants might have been impure, which would result in incorrect theoretical yield calculations. The student might have used poor lab techniques when conducting the lab, which can result in low percentage yield.

15. Cu_2S

16. $C_{20}H_{12}O_4$

17. 16.0 g

18. Sample answer:



19. Each atom of C has 12 times more mass than each H atom.

20. Sample answer: Ratios of the two elements

21. $C_{14}H_{18}N_2O_5$

22. The excess reactant will not limit or affect the reaction. Sample answer: Suppose you have one car and five wheels. You have an excess of wheels, but you have enough wheels for the car to move. The ability of the car to move is not affected by the fifth wheel.

23. 26.7%

24. Sample answer:
(1) Determine the mass of the beaker and of the piece of zinc.
(2) Add concentrated hydrochloric acid (excess) to the zinc under a fume hood, until all the zinc has reacted to form zinc chloride ($ZnCl_2$).
(3) Gently heat to evaporate unreacted hydrochloric acid and water.
(4) Determine the mass of zinc chloride that remains.

25. The idea is feasible.