CHAPTER 1

CHEMICAL FOUNDATIONS

Questions

11. A law summarizes what happens, e.g., law of conservation of mass in a chemical reaction or the ideal gas law, $PV = nRT$. A theory (model) is an attempt to explain why something happens. Dalton's atomic theory explains why mass is conserved in a chemical reaction. The kinetic molecular theory explains why pressure and volume are inversely related at constant temperature and moles of gas present, as well as explaining the other mathematical relationships summarized in $PV = nRT$.

12. a. At 8 a.m., approximately 57 cars pass through the intersection per hour.

    b. At 12 a.m. (midnight), only 1 or 2 cars pass through the intersection per hour.

    c. Traffic at the intersection is limited to less than 10 cars per hour from 8 p.m. to 5 a.m. Starting at 6 a.m., there is a steady increase in traffic through the intersection, peaking at 8 a.m. when approximately 57 cars pass per hour. Past 8 a.m. traffic moderates to about 40 cars through the intersection per hour until noon, and then decreases to 21 cars per hour by 3 p.m. Past 3 p.m. traffic steadily increases to a peak of 52 cars per hour at 5 p.m., and then steadily decreases to the overnight level of less than 10 cars through the intersection per hour.

    d. The traffic pattern through the intersection is directly related to the work schedules of the general population as well as to the store hours of the businesses in downtown.

    e. Run the same experiment on a Sunday, when most of the general population doesn't work and when a significant number of downtown stores are closed in the morning.

13. The fundamental steps are

    (1) making observations;
    (2) formulating hypotheses;
    (3) performing experiments to test the hypotheses.

    The key to the scientific method is performing experiments to test hypotheses. If after the test of time the hypotheses seem to account satisfactorily for some aspect of natural behavior, then the set of tested hypotheses turns into a theory (model). However, scientists continue to perform experiments to refine or replace existing theories. Hence, science is a dynamic or active process, not a static one.

14. A compound will always contain the same numbers (and types) of atoms. A given amount of hydrogen will react only with a specific amount of oxygen. Any excess oxygen will remain unreacted.

15. Law of conservation of mass: Mass is neither created nor destroyed. The total mass before a chemical reaction always equals the total mass after a chemical reaction.

Law of definite proportion: A given compound always contains exactly the same proportion of elements by mass. For example, water is always 1 g H for every 8 g oxygen.

Law of multiple proportions: When two elements form a series of compounds, the ratios of the mass of the second element that combine with 1 g of the first element always can be reduced to small whole numbers. For CO_2 and CO discussed in Section 1.4, the mass ratios of oxygen that react with 1 g carbon in each compound are in a 2 : 1 ratio.

16. Yes, 1.0 g H would react with 37.0 g ^{37}Cl, and 1.0 g H would react with 35.0 g ^{35}Cl.

No, the mass ratio of H/Cl would always be 1 g H/37 g Cl for ^{37}Cl and 1 g H/35 g Cl for ^{35}Cl. As long as we had pure ^{37}Cl or pure ^{35}Cl, the ratios will always hold. If we have a mixture (such as the natural abundance of chlorine), the ratio will also be constant as long as the composition of the mixture of the two isotopes does not change.

17. Natural niacin and commercially produced niacin have the exact same formula of $C_6H_5NO_2$. Therefore, both sources produce niacin having an identical nutritional value. There may be other compounds present in natural niacin that would increase the nutritional value, but the nutritional value due to just niacin is identical to the commercially produced niacin.

18. a. The smaller parts are electrons and the nucleus. The nucleus is broken down into protons and neutrons, which can be broken down into quarks. For our purpose, electrons, neutrons, and protons are the key smaller parts of an atom.

b. All atoms of hydrogen have 1 proton in the nucleus. Different isotopes of hydrogen have 0, 1, or 2 neutrons in the nucleus. Because we are talking about atoms, this implies a neutral charge, which dictates 1 electron present for all hydrogen atoms. If charged ions were included, then different ions/atoms of H could have different numbers of electrons.

c. Hydrogen atoms always have 1 proton in the nucleus, and helium atoms always have 2 protons in the nucleus. The number of neutrons can be the same for a hydrogen atom and a helium atom. Tritium (^3H) and ^4He both have 2 neutrons. Assuming neutral atoms, then the number of electrons will be 1 for hydrogen and 2 for helium.

d. Water (H_2O) is always 1 g hydrogen for every 8 g of O present, whereas H_2O_2 is always 1 g hydrogen for every 16 g of O present. These are distinctly different compounds, each with its own unique relative number and types of atoms present.

e. A chemical equation involves a reorganization of the atoms. Bonds are broken between atoms in the reactants, and new bonds are formed in the products. The number and types of atoms between reactants and products do not change. Because atoms are conserved in a chemical reaction, mass is also conserved.

19. J. J. Thomson's study of cathode-ray tubes led him to postulate the existence of negatively charged particles that we now call electrons. Ernest Rutherford and his alpha bombardment of metal foil experiments led him to postulate the nuclear atom–an atom with a tiny dense center of positive charge (the nucleus) with electrons moving about the nucleus at relatively large distances away; the distance is so large that an atom is mostly empty space.

20. The atom is composed of a tiny dense nucleus containing most of the mass of the atom. The nucleus itself is composed of neutrons and protons. Neutrons have a mass slightly larger than that of a proton and have no charge. Protons, on the other hand, have a 1+ relative charge as compared to the 1– charged electrons; the electrons move about the nucleus at relatively large distances. The volume of space that the electrons move about is so large, as compared to the nucleus, that we say an atom is mostly empty space.

21. The number and arrangement of electrons in an atom determine how the atom will react with other atoms. The electrons determine the chemical properties of an atom. The number of neutrons present determines the isotope identity.

22. Density = mass/volume; if the volumes are assumed equal, then the much more massive proton would have a much larger density than the relatively light electron.

23. For lighter, stable isotopes, the number of protons in the nucleus is about equal to the number of neutrons. When the number of protons and neutrons is equal to each other, the mass number (protons + neutrons) will be twice the atomic number (protons). Therefore, for lighter isotopes, the ratio of the mass number to the atomic number is close to 2. For example, consider ^{28}Si, which has 14 protons and (28 – 14 =) 14 neutrons. Here, the mass number to atomic number ratio is 28/14 = 2.0. For heavier isotopes, there are more neutrons than protons in the nucleus. Therefore, the ratio of the mass number to the atomic number increases steadily upward from 2 as the isotopes get heavier and heavier. For example, ^{238}U has 92 protons and (238 – 92 =) 146 neutrons. The ratio of the mass number to the atomic number for ^{238}U is 238/92 = 2.6.

24. Some elements exist as molecular substances. That is, hydrogen normally exists as H_2 molecules, not single hydrogen atoms. The same is true for N_2, O_2, F_2, Cl_2, Br_2, and I_2.

Exercises

Development of the Atomic Theory

25. a. The composition of a substance depends on the numbers of atoms of each element making up the compound (depends on the formula of the compound) and not on the composition of the mixture from which it was formed.

    b. Avogadro's hypothesis (law) implies that volume ratios are equal to molecule ratios at constant temperature and pressure. $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$. From the balanced equation, the volume of HCl produced will be twice the volume of H_2 (or Cl_2) reacted.

26. Avogadro's hypothesis (law) implies that volume ratios are equal to molecule ratios at constant temperature and pressure. Here, 1 volume of N_2 reacts with 3 volumes of H_2 to produce 2 volumes of the gaseous product or in terms of molecule ratios:

    $1\ N_2 + 3\ H_2 \rightarrow 2$ product

    In order for the equation to be balanced, the product must be NH_3.

27. From the law of definite proportions, a given compound always contains exactly the same proportion of elements by mass. The first sample of chloroform has a total mass of 12.0 g C + 106.4 g Cl + 1.01 g H = 119.41 g (carrying extra significant figures). The mass percent of carbon in this sample of chloroform is:

$$\frac{12.0 \text{ g C}}{119.41 \text{ g total}} \times 100 = 10.05\% \text{ C by mass}$$

From the law of definite proportions, the second sample of chloroform must also contain 10.05% C by mass. Let x = mass of chloroform in the second sample:

$$\frac{30.0 \text{ g C}}{x} \times 100 = 10.05, \quad x = 299 \text{ g chloroform}$$

28. A compound will always have a constant composition by mass. From the initial data given, the mass ratio of H : S : O in sulfuric acid (H_2SO_4) is:

$$\frac{2.02}{2.02} : \frac{32.07}{2.02} : \frac{64.00}{2.02} = 1 : 15.9 : 31.7$$

If we have 7.27 g H, then we will have 7.27 × 15.9 = 116 g S and 7.27 × 31.7 = 230. g O in the second sample of H_2SO_4.

29. Hydrazine: 1.44×10^{-1} g H/g N; ammonia: 2.16×10^{-1} g H/g N; hydrogen azide: 2.40×10^{-2} g H/g N. Let's try all of the ratios:

$$\frac{0.144}{0.0240} = 6.00; \quad \frac{0.216}{0.0240} = 9.00; \quad \frac{0.0240}{0.0240} = 1.00; \quad \frac{0.216}{0.144} = 1.50 = \frac{3}{2}$$

All the masses of hydrogen in these three compounds can be expressed as simple whole-number ratios. The g H/g N in hydrazine, ammonia, and hydrogen azide are in the ratios 6 : 9 : 1.

30. The law of multiple proportions does not involve looking at the ratio of the mass of one element with the total mass of the compounds. To illustrate the law of multiple proportions, we compare the mass of carbon that combines with 1.0 g of oxygen in each compound:

compound 1: 27.2 g C and 72.8 g O (100.0 – 27.2 = mass O)

compound 2: 42.9 g C and 57.1 g O (100.0 – 42.9 = mass O)

The mass of carbon that combines with 1.0 g of oxygen is:

compound 1: $\frac{27.2 \text{ g C}}{72.8 \text{ g O}} = 0.374$ g C/g O

compound 2: $\frac{42.9 \text{ g C}}{57.1 \text{ g O}} = 0.751$ g C/g O

$\frac{0.751}{0.374} = \frac{2}{1}$; this supports the law of multiple proportions because this carbon ratio is a whole number.

31. For CO and CO_2, it is easiest to concentrate on the mass of oxygen that combines with 1 g of carbon. From the formulas (two oxygen atoms per carbon atom in CO_2 versus one oxygen atom per carbon atom in CO), CO_2 will have twice the mass of oxygen that combines per gram of carbon as compared to CO. For CO_2 and C_3O_2, it is easiest to concentrate on the mass of carbon that combines with 1 g of oxygen. From the formulas (three carbon atoms per two oxygen atoms in C_3O_2 versus one carbon atom per two oxygen atoms in CO_2), C_3O_2 will have three times the mass of carbon that combines per gram of oxygen as compared to CO_2. As expected, the mass ratios are whole numbers as predicted by the law of multiple proportions.

32. Compound I: $\frac{14.0 \text{ g R}}{3.00 \text{ g Q}} = \frac{4.67 \text{ g R}}{1.00 \text{ g Q}}$; compound II: $\frac{7.00 \text{ g R}}{4.50 \text{ g Q}} = \frac{1.56 \text{ g R}}{1.00 \text{ g Q}}$

The ratio of the masses of R that combine with 1.00 g Q is: $\frac{4.67}{1.56} = 2.99 \approx 3$

As expected from the law of multiple proportions, this ratio is a small whole number.

Because compound I contains three times the mass of R per gram of Q as compared with compound II (RQ), the formula of compound I should be R_3Q.

33. Mass is conserved in a chemical reaction because atoms are conserved. Chemical reactions involve the reorganization of atoms, so formulas change in a chemical reaction, but the number and types of atoms do not change. Because the atoms do not change in a chemical reaction, mass must not change. In this equation we have two oxygen atoms and four hydrogen atoms both before and after the reaction occurs.

34. Mass is conserved in a chemical reaction.

	ethanol +	oxygen $\rightarrow$	water +	carbon dioxide
Mass:	46.0 g	96.0 g	54.0 g	?

Mass of reactants = 46.0 + 96.0 = 142.0 g = mass of products

142.0 g = 54.0 g + mass of CO_2, mass of CO_2 = 142.0 – 54.0 = 88.0 g

35. To get the atomic mass of H to be 1.00, we divide the mass of hydrogen that reacts with 1.00 g of oxygen by 0.126; that is, $\frac{0.126}{0.126} = 1.00$. To get Na, Mg, and O on the same scale, we do the same division.

Na: $\frac{2.875}{0.126} = 22.8$; Mg: $\frac{1.500}{0.126} = 11.9$; O: $\frac{1.00}{0.126} = 7.94$

	H	O	Na	Mg
Relative value	1.00	7.94	22.8	11.9
Accepted value	1.008	16.00	22.99	24.31

For your information, the atomic masses of O and Mg are incorrect. The atomic masses of H and Na are close to the values given in the periodic table. Something must be wrong about the assumed formulas of the compounds. It turns out the correct formulas are H_2O, Na_2O, and MgO. The smaller discrepancies result from the error in the assumed atomic mass of H.

36. If the formula is InO, then one atomic mass of In would combine with one atomic mass of O, or:

$$\frac{A}{16.00} = \frac{4.784 \text{ g In}}{1.000 \text{ g O}}, \text{ A = atomic mass of In = 76.54}$$

If the formula is In_2O_3, then two times the atomic mass of In will combine with three times the atomic mass of O, or:

$$\frac{2A}{(3)16.00} = \frac{4.784 \text{ g In}}{1.000 \text{ g O}}, \text{ A = atomic mass of In = 114.8}$$

The latter number is the atomic mass of In used in the modern periodic table.

The Nature of the Atom

37. From section 1-7, the nucleus has "a diameter of about 10^{-13} cm" and the electrons "move about the nucleus at an average distance of about 10^{-8} cm from it." We will use these statements to help determine the densities. Density of hydrogen nucleus (contains one proton only):

$$V_{nucleus} = \frac{4}{3}\pi r^3 = \frac{4}{3}(3.14)(5 \times 10^{-14} \text{ cm})^3 = 5 \times 10^{-40} \text{ cm}^3$$

$$d = \text{density} = \frac{1.67 \times 10^{-24} \text{ g}}{5 \times 10^{-40} \text{ cm}^3} = 3 \times 10^{15} \text{ g/cm}^3$$

Density of H atom (contains one proton and one electron):

$$V_{atom} = \frac{4}{3}(3.14)(1 \times 10^{-8} \text{ cm})^3 = 4 \times 10^{-24} \text{ cm}^3$$

$$d = \frac{1.67 \times 10^{-24} \text{ g} + 9 \times 10^{-28} \text{g}}{4 \times 10^{-24} \text{ cm}^3} = 0.4 \text{ g/cm}^3$$

38. Because electrons move about the nucleus at an average distance of about 1×10^{-8} cm, the diameter of an atom will be about 2×10^{-8} cm. Let's set up a ratio:

$$\frac{\text{diameter of nucleus}}{\text{diameter of atom}} = \frac{1 \text{ mm}}{\text{diameter of model}} = \frac{1 \times 10^{-13} \text{ cm}}{2 \times 10^{-8} \text{ cm}}; \text{ solving:}$$

diameter of model = 2×10^5 mm = 200 m

39. $5.93 \times 10^{-18} \text{ C} \times \frac{1 \text{ electron charge}}{1.602 \times 10^{-19} \text{ C}}$ = 37 negative (electron) charges on the oil drop

40. First, divide all charges by the smallest quantity, 6.40×10^{-13}:

$$\frac{2.56 \times 10^{-12}}{6.40 \times 10^{-13}} = 4.00; \quad \frac{7.68}{0.640} = 12.0; \quad \frac{3.84}{0.640} = 6.00$$

Because all charges are whole-number multiples of 6.40×10^{-13} zirkombs, the charge on one electron could be 6.40×10^{-13} zirkombs. However, 6.40×10^{-13} zirkombs could be the charge of two electrons (or three electrons, etc.). All one can conclude is that the charge of an electron is 6.40×10^{-13} zirkombs or an integer fraction of 6.40×10^{-13} zirkombs.

41. Z is the atomic number and is equal to the number of protons in the nucleus. A is the mass number and is equal to the number of protons plus neutrons in the nucleus. X is the symbol of the element. See the front cover of the text which has a listing of the symbols for the various elements and corresponding atomic number or see the periodic table on the cover to determine the identity of the various atoms. Because all of the atoms have equal numbers of protons and electrons, each atom is neutral in charge.

a. ${}^{23}_{11}Na$ b. ${}^{19}_{9}F$ c. ${}^{16}_{8}O$

42. The atomic number for carbon is 6. ${}^{14}C$ has 6 protons, 14 – 6 = 8 neutrons, and 6 electrons in the neutral atom. ${}^{12}C$ has 6 protons, 12 – 6 = 6 neutrons, and 6 electrons in the neutral atom. The only difference between an atom of ${}^{14}C$ and an atom of ${}^{12}C$ is that ${}^{14}C$ has two additional neutrons.

43. a. ${}^{79}_{35}Br$: 35 protons, 79 – 35 = 44 neutrons. Because the charge of the atom is neutral, the number of protons = the number of electrons = 35.

b. ${}^{81}_{35}Br$: 35 protons, 46 neutrons, 35 electrons

c. ${}^{239}_{94}Pu$: 94 protons, 145 neutrons, 94 electrons

d. ${}^{133}_{55}Cs$: 55 protons, 78 neutrons, 55 electrons

e. ${}^{3}_{1}H$: 1 proton, 2 neutrons, 1 electron

f. ${}^{56}_{26}Fe$: 26 protons, 30 neutrons, 26 electrons

44. a. ${}^{235}_{92}U$: 92 p, 143 n, 92 e b. ${}^{27}_{13}Al$: 13 p, 14 n, 13 e c. ${}^{57}_{26}Fe$: 26 p, 31 n, 26 e

d. ${}^{208}_{82}Pb$: 82 p, 126 n, 82 e e. ${}^{86}_{37}Rb$: 37 p, 49 n, 37 e f. ${}^{41}_{20}Ca$: 20 p, 21 n, 20 e

45. a. Element 8 is oxygen. A = mass number = 9 + 8 = 17; ${}^{17}_{8}O$

b. Chlorine is element 17. ${}^{37}_{17}Cl$ c. Cobalt is element 27. ${}^{60}_{27}Co$

d. Z = 26; A = 26 + 31 = 57; ${}^{57}_{26}Fe$ e. Iodine is element 53. ${}^{131}_{53}I$

f. Lithium is element 3. ${}^{7}_{3}Li$

46. a. Cobalt is element 27. A = mass number = 27 + 31 = 58; ${}^{58}_{27}\text{Co}$

b. ${}^{10}_{5}\text{B}$ c. ${}^{23}_{12}\text{Mg}$ d. ${}^{132}_{53}\text{I}$ e. ${}^{47}_{20}\text{Ca}$ f. ${}^{65}_{29}\text{Cu}$

47. a. Ba is element 56. Ba^{2+} has 56 protons, so Ba^{2+} must have 54 electrons in order to have a net charge of 2+.

b. Zn is element 30. Zn^{2+} has 30 protons and 28 electrons.

c. N is element 7. N^{3-} has 7 protons and 10 electrons.

d. Rb is element 37, Rb^{+} has 37 protons and 36 electrons.

e. Co is element 27. Co^{3+} has 27 protons and 24 electrons.

f. Te is element 52. Te^{2-} has 52 protons and 54 electrons.

g. Br is element 35. Br^{-} has 35 protons and 36 electrons.

48. a. ${}^{24}_{12}\text{Mg}$: 12 protons, 12 neutrons, 12 electrons

b. ${}^{24}_{12}\text{Mg}^{2+}$: 12 p, 12 n, 10 e

c. ${}^{59}_{27}\text{Co}^{2+}$: 27 p, 32 n, 25 e

d. ${}^{59}_{27}\text{Co}^{3+}$: 27 p, 32 n, 24 e

e. ${}^{59}_{27}\text{Co}$: 27 p, 32 n, 27 e

f. ${}^{79}_{34}\text{Se}$: 34 p, 45 n, 34 e

g. ${}^{79}_{34}\text{Se}^{2-}$: 34 p, 45 n, 36 e

h. ${}^{63}_{28}\text{Ni}$: 28 p, 35 n, 28 e

i. ${}^{59}_{28}\text{Ni}^{2+}$: 28 p, 31 n, 26 e

49. Atomic number = 63 (Eu); net charge = +63 – 60 = 3+; mass number = 63 + 88 = 151; symbol: ${}^{151}_{63}\text{Eu}^{3+}$

Atomic number = 50 (Sn); mass number = 50 + 68 = 118; net charge = +50 – 48 = 2+; symbol: ${}^{118}_{50}\text{Sn}^{2+}$

50. Atomic number = 16 (S); net charge = +16 – 18 = 2–; mass number = 16 + 18 = 34; symbol: ${}^{34}_{16}\text{S}^{2-}$

Atomic number = 16 (S); net charge = +16 – 18 = 2–; mass number = 16 + 16 = 32; symbol: ${}^{32}_{16}\text{S}^{2-}$

51.

Symbol	Number of protons in nucleus	Number of neutrons in nucleus	Number of electrons	Net charge
$^{238}_{92}U$	92	146	92	0
$^{40}_{20}Ca^{2+}$	20	20	18	2+
$^{51}_{23}V^{3+}$	23	28	20	3+
$^{89}_{39}Y$	39	50	39	0
$^{79}_{35}Br^{-}$	35	44	36	1−
$^{31}_{15}P^{3-}$	15	16	18	3−

52.

Symbol	Number of protons in nucleus	Number of neutrons in nucleus	Number of electrons	Net charge
$^{53}_{26}Fe^{2+}$	26	27	24	2+
$^{59}_{26}Fe^{3+}$	26	33	23	3+
$^{210}_{85}At^{-}$	85	125	86	1–
$^{27}_{13}Al^{3+}$	13	14	10	3+
$^{128}_{52}Te^{2-}$	52	76	54	2–

Additional Exercises

53. $^{53}_{26}Fe^{2+}$ has 26 protons, 53 – 26 = 27 neutrons, and two fewer electrons than protons (24 electrons) in order to have a net charge of 2+.

54. a. False. Neutrons have no charge; therefore, all particles in a nucleus are not charged.

    b. False. The atom is best described as having a tiny dense nucleus containing most of the mass of the atom with the electrons moving about the nucleus at relatively large distances away; so much so that an atom is mostly empty space.

    c. False. The mass of the nucleus makes up most of the mass of the entire atom.

    d. True.

    e. False. The number of protons in a neutral atom must equal the number of electrons.

55. a. Element 15 is phosphorus, P. This atom has 15 protons and 31–15 = 16 neutrons.

    b. Element 53 is iodine, I. 53 protons; 74 neutrons

    c. Element 19 is potassium, K. 19 protons; 20 neutrons

    d. Element 70 is ytterbium, Yb. 70 protons; 103 neutrons

56. For a 2– charge having 36 electrons, X^{2-} must have two fewer protons than electrons or 34 protons. This isotope of X has 79 – 34 = 45 neutrons, and from the periodic table, X is selenium.

    a. True. X^{2-} has 36 electrons and 34 protons.

    b. False. The isotope has 34 protons.

    c. False. The isotope has 45 neutrons.

    d. False. The identity is selenium, Se.

57. Because the charge of the ion is 2+, this ion has two more protons than electrons. Therefore, the element has 88 protons, which identifies it as radium, Ra. 230 – 88 = 142 neutrons.

58. The alchemists were incorrect. The solid residue must have come from the flask.

59. Mass is conserved in a chemical reaction.

	chromium(III) oxide	+	aluminum	→	chromium	+	aluminum oxide
Mass:	34.0 g		12.1 g		23.3 g		?

Mass of aluminum oxide produced = (34.0 + 12.1) – 23.3 = 22.8 g

ChemWork Problems

60.

Number of protons in nucleus	Number of neutrons in nucleus	Symbol
9	10	$^{19}_{9}F$
13	14	$^{27}_{13}Al$
53	74	$^{127}_{53}I$
34	45	$^{79}_{34}Se$
16	16	$^{32}_{16}S$

61.

Symbol	Number of protons in nucleus	Number of neutrons in nucleus
$^{4}_{2}He$	2	2
$^{20}_{10}Ne$	10	10
$^{48}_{22}Ti$	22	26
$^{190}_{76}Os$	76	114
$^{50}_{27}Co$	27	23

62.

Symbol	Number of protons in nucleus	Number of neutrons in nucleus	Number of electrons
$^{120}_{50}Sn$	50	70	50
$^{25}_{12}Mg^{2+}$	12	13	10
$^{56}_{26}Fe^{2+}$	26	30	24
$^{79}_{34}Se$	34	45	34
$^{35}_{17}Cl$	17	18	17
$^{63}_{29}Cu$	29	34	29

63. a. True

b. False; this was J. J. Thomson.

c. False; a proton is about 1800 times more massive than an electron.

d. The nucleus contains the protons and the neutrons.

Challenge Problems

64. a. One possibility is that rope B is not attached to anything and rope A and rope C are connected via a pair of pulleys and/or gears.

b. Try to pull rope B out of the box. Measure the distance moved by C for a given movement of A. Hold either A or C firmly while pulling on the other rope.

65. a. Both compounds have C_2H_6O as the formula. Because they have the same formula, their mass percent composition will be identical. However, these are different compounds with different properties because the atoms are bonded together differently. These compounds are called isomers of each other.

b. When wood burns, most of the solid material in wood is converted to gases, which escape. The gases produced are most likely CO_2 and H_2O.

c. The atom is not an indivisible particle but is instead composed of other smaller particles, called electrons, neutrons, and protons.

d. The two hydride samples contain different isotopes of either hydrogen and/or lithium. Although the compounds are composed of different isotopes, their properties are similar because different isotopes of the same element have similar properties (except, of course, their mass).

66. Because the gases are at the same temperature and pressure, the volumes are directly proportional to the number of molecules present. Let's consider hydrogen and oxygen to be monatomic gases and that water has the simplest possible formula (HO). We have the equation:

$$H + O \rightarrow HO$$

But the volume ratios are also equal to the molecule ratios, which correspond to the coefficients in the equation:

$$2\ H + O \rightarrow 2\ HO$$

Because atoms cannot be created nor destroyed in a chemical reaction, this is not possible. To correct this, we can make oxygen a diatomic molecule:

$$2\ H + O_2 \rightarrow 2\ HO$$

This does not require hydrogen to be diatomic. Of course, if we know water has the formula H_2O, we get:

$$2\ H + O_2 \rightarrow 2\ H_2O$$

The only way to balance this is to make hydrogen diatomic:

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

67. Avogadro proposed that equal volumes of gases (at constant temperature and pressure) contain equal numbers of molecules. In terms of balanced equations, Avogadro's hypothesis (law) implies that volume ratios will be identical to molecule ratios. Assuming one molecule of octane reacting, then 1 molecule of C_xH_y produces 8 molecules of CO_2 and 9 molecules of H_2O. $C_xH_y + n\ O_2 \rightarrow 8\ CO_2 + 9\ H_2O$. Because all the carbon in octane ends up as carbon in CO_2, octane must contain 8 atoms of C. Similarly, all hydrogen in octane ends up as hydrogen in H_2O, so one molecule of octane must contain $9 \times 2 = 18$ atoms of H. Octane formula = C_8H_{18}, and the ratio of C : H = 8 : 18 or 4 : 9.

68. From Section 1-7 of the text, the average diameter of the nucleus is about 10^{-13} cm, and the electrons move about the nucleus at an average distance of about 10^{-8} cm. From this, the diameter of an atom is about 2×10^{-8} cm.

$$\frac{2 \times 10^{-8}\ \text{cm}}{1 \times 10^{-13}\ \text{cm}} = 2 \times 10^5; \quad \frac{1\ \text{mi}}{1\ \text{grape}} = \frac{5280\ \text{ft}}{1\ \text{grape}} = \frac{63{,}360\ \text{in}}{1\ \text{grape}}$$

Because the grape needs to be 2×10^5 times smaller than a mile, the diameter of the grape would need to be $63{,}360/(2 \times 10^5) \approx 0.3$ in. This is a reasonable size for a small grape.

69. Let X_a be the formula for the atom/molecule X, Y_b be the formula for the atom/molecule Y, X_cY_d be the formula of compound I between X and Y, and X_eY_f be the formula of compound

II between X and Y. Using the volume data, the following would be the balanced equations for the production of the two compounds.

$X_a + 2\ Y_b \rightarrow 2\ X_cY_d$; $2\ X_a + Y_b \rightarrow 2\ X_eY_f$

From the balanced equations, $a = 2c = e$ and $b = d = 2f$.

Substituting into the balanced equations:

$X_{2c} + 2\ Y_{2f} \rightarrow 2\ X_cY_{2f}$
$2\ X_{2c} + Y_{2f} \rightarrow 2\ X_{2c}Y_f$

For simplest formulas, assume that $c = f = 1$. Thus:

$X_2 + 2\ Y_2 \rightarrow 2\ XY_2$ and $2\ X_2 + Y_2 \rightarrow 2\ X_2Y$

Compound I = XY_2: If X has relative mass of 1.00, $\dfrac{1.00}{1.00 + 2y} = 0.3043$, $y = 1.14$.

Compound II = X_2Y: If X has relative mass of 1.00, $\dfrac{2.00}{2.00 + y} = 0.6364$, $y = 1.14$.

The relative mass of Y is 1.14 times that of X. Thus, if X has an atomic mass of 100, then Y will have an atomic mass of 114.

Integrative Exercises

70. Number of electrons in the unknown ion:

$$2.55 \times 10^{-26}\ \text{g} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \frac{1\ \text{electron}}{9.11 \times 10^{-31}\ \text{kg}} = 28\ \text{electrons}$$

Number of protons in the unknown ion:

$$5.34 \times 10^{-23}\ \text{g} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \frac{1\ \text{proton}}{1.67 \times 10^{-27}\ \text{kg}} = 32\ \text{protons}$$

Therefore, this ion has 32 protons and 28 electrons. This is element number 32, germanium (Ge). The net charge is 4+ because four electrons have been lost from a neutral germanium atom.

The number of electrons in the unknown atom:

$$3.92 \times 10^{-26}\ \text{g} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \frac{1\ \text{electron}}{9.11 \times 0^{-31}\ \text{kg}} = 43\ \text{electrons}$$

In a neutral atom, the number of protons and electrons is the same. Therefore, this is element 43, technetium (Tc).

The number of neutrons in the technetium atom:

$$9.35 \times 10^{-23}\ \text{g} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \frac{1\ \text{proton}}{1.67 \times 10^{-27}\ \text{kg}} = 56\ \text{neutrons}$$

The mass number is the sum of the protons and neutrons. In this atom, the mass number is 43 protons + 56 neutrons = 99. Thus this atom and its mass number is ^{99}Tc.

71. Most of the mass of the atom is due to the protons and the neutrons in the nucleus, and protons and neutrons have about the same mass (1.67×10^{-24} g). The ratio of the mass of the molecule to the mass of a nuclear particle will give a good approximation of the number of nuclear particles (protons and neutrons) present.

$$\frac{7.31 \times 10^{-23}\ \text{g}}{1.67 \times 10^{-24}\ \text{g}} = 43.8 \approx 44\ \text{nuclear particles}$$

Thus there are 44 protons and neutrons present. If the number of protons equals the number of neutrons, we have 22 protons in the molecule. One possibility would be the molecule CO_2 [6 + 2(8) = 22 protons].

Marathon Problem

72. a. For each set of data, divide the larger number by the smaller number to determine relative masses.

$$\frac{0.602}{0.295} = 2.04; \quad A = 2.04 \text{ when } B = 1.00$$

$$\frac{0.401}{0.172} = 2.33; \quad C = 2.33 \text{ when } B = 1.00$$

$$\frac{0.374}{0.320} = 1.17; \quad C = 1.17 \text{ when } A = 1.00$$

To have whole numbers, multiply the results by 3.

Data set 1: A = 6.1 and B = 3.0

Data set 2: C = 7.0 and B = 3.0

Data set 3: C = 3.5 and A = 3.0 or C = 7.0 and A = 6.0

Assuming 6.0 for the relative mass of A, the relative masses would be A = 6.0, B = 3.0, and C = 7.0 (if simplest formulas are assumed).

b. Gas volumes are proportional to the number of molecules present. There are many possible correct answers for the balanced equations. One such solution that fits the gas volume data is:

$$6\ A_2 + B_4 \rightarrow 4\ A_3B$$

$$B_4 + 4\ C_3 \rightarrow 4\ BC_3$$

$$3\ A_2 + 2\ C_3 \rightarrow 6\ AC$$

In any correct set of reactions, the calculated mass data must match the mass data given initially in the problem. Here, the new table of relative masses would be:

$$\frac{6\,(\text{mass } A_2)}{\text{mass } B_4} = \frac{0.602}{0.295};\ \text{mass } A_2 = 0.340(\text{mass } B_4)$$

$$\frac{4\,(\text{mass } C_3)}{\text{mass } B_4} = \frac{0.401}{0.172};\ \text{mass } C_3 = 0.583(\text{mass } B_4)$$

$$\frac{2\,(\text{mass } C_3)}{3\,(\text{mass } A_2)} = \frac{0.374}{0.320};\ \text{mass } A_2 = 0.570(\text{mass } C_3)$$

Assume some relative mass number for any of the masses. We will assume that mass B = 3.0, so mass B_4 = 4(3.0) = 12.

Mass C_3 = 0.583(12) = 7.0, mass C = 7.0/3

Mass A_2 = 0.570(7.0) = 4.0, mass A = 4.0/2 = 2.0

When we assume a relative mass for B = 3.0, then A = 2.0 and C = 7.0/3. The relative masses having all whole numbers would be A = 6.0, B = 9.0, and C = 7.0.

Note that any set of balanced reactions that confirms the initial mass data is correct. This is just one possibility.