CHAPTER 2

ATOMS, MOLECULES, AND IONS

Development of the Atomic Theory

18. 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 hydrogen 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 2.2, the mass ratios of oxygen that react with 1 g carbon in each compound are in a 2 : 1 ratio.

19. From Avogadro's hypothesis (law), volume ratios are equal to molecule ratios at constant temperature and pressure. Therefore, we can write a balanced equation using the volume data, $Cl_2 + 5\ F_2 \rightarrow 2\ X$. Two molecules of X contain 10 atoms of F and two atoms of Cl. The formula of X is ClF_5 for a balanced equation.

20. 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 proportional to molecule ratios at constant temperature and pressure. $H_2 + Cl_2 \rightarrow 2\ HCl$. From the balanced equation, the volume of HCl produced will be twice the volume of H_2 (or Cl_2) reacted.

21. 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.

22. 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.

23. 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.

24. Compound 1: 21.8 g C and 58.2 g O (80.0 – 21.8 = mass O)

Compound 2: 34.3 g C and 45.7 g O (80.0 – 34.3 = mass O)

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

Compound 1: $\frac{21.8 \text{ g C}}{58.2 \text{ g O}} = 0.375$ g C/g O

Compound 2: $\frac{34.3 \text{ g C}}{45.7 \text{ g O}} = 0.751$ g C/g O

The ratio of the masses of carbon that combine with 1 g of oxygen is $\frac{0.751}{0.375} = \frac{2}{1}$; this supports the law of multiple proportions because this carbon ratio is a small whole number.

25. To get the atomic mass of H to be 1.00, we divide the mass that reacts with 1.00 g of oxygen by 0.126, that is, 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	1.0079	15.999	22.99	24.31

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

The Nature of the Atom

26. Deflection of cathode rays by magnetic and electrical fields led to the conclusion that they were negatively charged. The cathode ray was produced at the negative electrode and repelled by the negative pole of the applied electrical field. β particles are electrons. A cathode ray is a stream of electrons (β particles).

27. From section 2.6, 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):

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

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

28. From Section 2.6 of the text, the average diameter of the nucleus is approximately 10^{-13} cm, and the electrons move about the nucleus at an average distance of approximately 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.

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

$$\frac{2.56 \times 10^{-12}}{6.40 \times 10^{-13}} = 4.00; \quad \frac{7.68}{0.640} = 12.00; \quad \frac{3.84}{0.640} = 6.00$$

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}.

30. The proton and neutron have similar mass, with the mass of the neutron slightly larger than that of the proton. Each of these particles has a mass approximately 1800 times greater than that of an electron. The combination of the protons and the neutrons in the nucleus makes up the bulk of the mass of an atom, but the electrons make the greatest contribution to the chemical properties of the atom.

31. If the plum pudding model were correct (a diffuse positive charge with electrons scattered throughout), then α particles should have traveled through the thin foil with very minor deflections in their path. This was not the case because a few of the α particles were deflected at very large angles. Rutherford reasoned that the large deflections of these α particles could be caused only by a center of concentrated positive charge that contains most of the atom's mass (the nuclear model of the atom).

Elements, Ions, and the Periodic Table

32. a. A molecule has no overall charge (an equal number of electrons and protons are present). Ions, on the other hand, have electrons added to form anions (negatively charged ions) or electrons removed to form cations (positively charged ions).

    b. The sharing of electrons between atoms is a covalent bond. An ionic bond is the force of attraction between two oppositely charged ions.

    c. A molecule is a collection of atoms held together by covalent bonds. A compound is composed of two or more different elements having constant composition. Covalent and/or ionic bonds can hold the atoms together in a compound. Another difference is that molecules do not necessarily have to be compounds. H_2 is two hydrogen atoms held together by a covalent bond. H_2 is a molecule, but it is not a compound; H_2 is a diatomic element.

    d. An anion is a negatively charged ion, for example, Cl^-, O^{2-}, and SO_4^{2-} are all anions. A cation is a positively charged ion, for example, Na^+, Fe^{3+}, and NH_4^+ are all cations.

33. The atomic number of an element is equal to the number of protons in the nucleus of an atom of that element. The mass number is the sum of the number of protons plus neutrons in the nucleus. The atomic mass is the actual mass of a particular isotope (including electrons). As is discussed in Chapter 3, the average mass of an atom is taken from a measurement made on a large number of atoms. The average atomic mass value is listed in the periodic table.

34. a. Metals: Mg, Ti, Au, Bi, Ge, Eu, and Am. Nonmetals: Si, B, At, Rn, and Br.

    b. Si, Ge, B, and At. The elements at the boundary between the metals and the nonmetals are B, Si, Ge, As, Sb, Te, Po, and At. Aluminum has mostly properties of metals, so it is generally not classified as a metalloid.

35. a. The noble gases are He, Ne, Ar, Kr, Xe, and Rn (helium, neon, argon, krypton, xenon, and radon). Radon has only radioactive isotopes. In the periodic table, the whole number enclosed in parentheses is the mass number of the longest-lived isotope of the element.

    b. promethium (Pm) and technetium (Tc)

36. Carbon is a nonmetal. Silicon and germanium are called metalloids as they exhibit both metallic and nonmetallic properties. Tin and lead are metals. Thus metallic character increases as one goes down a family in the periodic table. The metallic character decreases from left to right across the periodic table.

37. Use the periodic table to identify the elements.

a. Cl; halogen
b. Be; alkaline earth metal
c. Eu; lanthanide metal
d. Hf; transition metal
e. He; noble gas
f. U; actinide metal
g. Cs; alkali metal

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

39. 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.

40. a. transition metals
b. alkaline earth metals
c. alkali metals
d. noble gases
e. halogens

41. 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

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

b. $^{10}_{5}B$
c. $^{23}_{12}Mg$
d. $^{132}_{53}I$
e. $^{19}_{9}F$
f. $^{65}_{29}Cu$

43. a. $^{24}_{12}Mg$: 12 protons, 12 neutrons, 12 electrons

b. $^{24}_{12}Mg^{2+}$: 12 p, 12 n, 10 e
c. $^{59}_{27}Co^{2+}$: 27 p, 32 n, 25 e
d. $^{59}_{27}Co^{3+}$: 27 p, 32 n, 24 e
e. $^{59}_{27}Co$: 27 p, 32 n, 27 e
f. $^{79}_{34}Se$: 34 p, 45 n, 34 e
g. $^{79}_{34}Se^{2-}$: 34 p, 45 n, 36 e
h. $^{63}_{28}Ni$: 28 p, 35 n, 28 e
i. $^{59}_{28}Ni^{2+}$: 28 p, 31 n, 26 e

44.

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–

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

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

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

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

47. In ionic compounds, metals lose electrons to form cations, and nonmetals gain electrons to form anions. Group 1A, 2A, and 3A metals form stable 1+, 2+, and 3+ charged cations, respectively. Group 5A, 6A, and 7A nonmetals form 3–, 2–, and 1– charged anions, respectively.

a. Lose 2 e^- to form Ra^{2+}. b. Lose 3 e^- to form In^{3+}. c. Gain 3 e^- to form P^{3-}.

d. Gain 2 e^- to form Te^{2-}. e. Gain 1 e^- to form Br^-. f. Lose 1 e^- to form Rb^+.

48. See Exercise 47 for a discussion of charges various elements form when in ionic compounds.

a. Element 13 is Al. Al forms 3+ charged ions in ionic compounds. Al^{3+}

b. Se^{2-} c. Ba^{2+} d. N^{3-} e. Fr^+ f. Br^-

Nomenclature

49. $AlCl_3$, aluminum chloride; $CrCl_3$, chromium(III) chloride; ICl_3, iodine trichloride; $AlCl_3$ and $CrCl_3$ are ionic compounds following the rules for naming ionic compounds. The major difference is that $CrCl_3$ contains a transition metal (Cr) that generally exhibits two or more stable charges when in ionic compounds. We need to indicate which charged ion we have in the compound. This is generally true whenever the metal in the ionic compound is a transition metal. ICl_3 is made from only nonmetals and is a covalent compound. Predicting formulas for covalent compounds is extremely difficult. Because of this, we need to indicate the number of each nonmetal in the binary covalent compound. The exception is when there is only one of the first species present in the formula; when this is the case, mono- is not used (it is assumed).

50. a. Dinitrogen monoxide is correct. N and O are both nonmetals resulting in a covalent compound. We need to use the covalent rules of nomenclature. The other two names are for ionic compounds.

    b. Copper(I) oxide is correct. With a metal in a compound, we have an ionic compound. Because copper, like most transition metals, forms at least a couple of different stable charged ions, we must indicate the charge on copper in the name. Copper oxide could be CuO or Cu_2O, hence why we must give the charge of most transition metal compounds. Dicopper monoxide is the name if this were a covalent compound, which it is not.

    c. Lithium oxide is correct. Lithium forms 1+ charged ions in stable ionic compounds. Because lithium is assumed to form 1+ ions in compounds, we do not need to indicate the charge of the metal ion in the compound. Dilithium monoxide would be the name if Li_2O were a covalent compound (a compound composed of only nonmetals).

51. a. mercury(I) oxide b. iron(III) bromide
    c. cobalt(II) sulfide d. titanium(IV) chloride
    e. Sn_3N_2 f. CoI_3
    g. HgO h. CrS_3

52. a. barium sulfite b. sodium nitrite
    c. potassium permanganate d. potassium dichromate
    e. $Cr(OH)_3$ f. $Mg(CN)_2$
    g. $Pb(CO_3)_2$ h. $NH_4C_2H_3O_2$

53. a. sulfur difluoride b. dinitrogen tetroxide
    c. iodine trichloride d. tetraphosphorus hexoxide

54. a. sodium perchlorate b. magnesium phosphate
    c. aluminum sulfate d. sulfur difluoride

e. sulfur hexafluoride
f. sodium hydrogen phosphate
g. sodium dihydrogen phosphate
h. lithium nitride
i. sodium hydroxide
j. magnesium hydroxide
k. aluminum hydroxide
l. silver chromate

55. a. copper(I) iodide
b. copper(II) iodide
c. cobalt(II) iodide
d. sodium carbonate
e. sodium hydrogen carbonate or sodium bicarbonate
f. tetrasulfur tetranitride
g. selenium tetrabromide
h. sodium hypochlorite
i. barium chromate
j. ammonium nitrate

56. a. acetic acid
b. ammonium nitrite
c. colbalt(III) sulfide
d. iodine monochloride
e. lead(II) phosphate
f. potassium chlorate
g. sulfuric acid
h. strontium nitride
i. aluminum sulfite
j. tin(IV) oxide
k. sodium chromate
l. hypochlorous acid

57. a. SO_2
b. SO_3
c. Na_2SO_3
d. $KHSO_3$
e. Li_3N
f. $Cr_2(CO_3)_3$
g. $Cr(C_2H_3O_2)_2$
h. SnF_4
i. NH_4HSO_4: composed of NH_4^+ and HSO_4^- ions
j. $(NH_4)_2HPO_4$
k. $KClO_4$
l. NaH
m. HBrO
n. HBr

58. a. Na_2O
b. Na_2O_2
c. KCN
d. $Cu(NO_3)_2$
e. $SiCl_4$
f. PbO
g. PbO_2
h. CuCl
i. GaAs: We would predict the stable ions to be Ga^{3+} and As^{3-}.
j. CdSe
k. ZnS
l. Hg_2Cl_2: Mercury(I) exists as Hg_2^{2+}.
m. HNO_2
n. P_2O_5

59. a. $Pb(C_2H_3O_2)_2$; lead(II) acetate
b. $CuSO_4$; copper(II) sulfate
c. CaO; calcium oxide
d. $MgSO_4$; magnesium sulfate
e. $Mg(OH)_2$; magnesium hydroxide
f. $CaSO_4$; calcium sulfate
g. N_2O; dinitrogen monoxide or nitrous oxide (common name)

60. a. Iron forms 2+ and 3+ charged ions; we need to include a Roman numeral for iron. Iron(III) chloride is correct.

b. This is a covalent compound so use the covalent rules. Nitrogen dioxide is correct.

c. This is an ionic compound, so use the ionic rules. Calcium oxide is correct. Calcium only forms stable 2+ ions when in ionic compounds, so no Roman numeral is needed.

d. This is an ionic compound, so use the ionic rules. Aluminum sulfide is correct.

e. This is an ionic compound, so use the ionic rules. Mg is magnesium. Magnesium acetate is correct.

f. Because phosphate has a 3– charge, the charge on iron is 3+. Iron(III) phosphate is correct.

g. This is a covalent compound, so use the covalent rules. Diphosphorus pentasulfide is correct.

h. Because each sodium is 1+ charged, we have the O_2^{2-} (peroxide) ion present. Sodium peroxide is correct. Note that sodium oxide would be Na_2O.

i. HNO_3 is nitric acid, not nitrate acid. Nitrate acid does not exist.

j. H_2S is hydrosulfuric acid or dihydrogen sulfide or just hydrogen sulfide (common name). H_2SO_4 is sulfuric acid.

61. a. nitric acid, HNO_3 b. perchloric acid, $HClO_4$ c. acetic acid, $HC_2H_3O_2$

d. sulfuric acid, H_2SO_4 e. phosphoric acid, H_3PO_4

Additional Exercises

62. 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.

63. a. This represents ionic bonding. Ionic bonding is the electrostatic attraction between anions and cations.

b. This represents covalent bonding where electrons are shared between two atoms. This could be the space-filling model for H_2O or SF_2 or NO_2, etc.

64. J. J. Thomson discovered electrons. He postulated that all atoms must contain electrons, but Thomson also postulated that atoms must contain positive charge in order for the atom to be electrically neutral. Henri Becquerel discovered radioactivity. Lord Rutherford proposed the nuclear model of the atom. Dalton's original model proposed that atoms were indivisible particles (that is, atoms had no internal structure). Thomson and Becquerel discovered subatomic particles, and Rutherford's model attempted to describe the internal structure of the atom composed of these subatomic particles. In addition, the existence of isotopes, atoms of the same element but with different mass, had to be included in the model.

65. The equation for the reaction between the elements of sodium and chlorine is $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$. The sodium reactant exists as singular sodium atoms packed together very tightly and in a very organized fashion. This type of packing of atoms represents the solid phase. The chlorine reactant exists as Cl_2 molecules. In the picture of chlorine, there is a lot of empty space present. This only occurs in the gaseous phase. When sodium and chlorine react, the ionic compound NaCl is the product. NaCl exists as separate Na^+ and Cl^- ions. Because the ions are packed very closely together and are packed in a very organized fashion, NaCl is depicted in the solid phase.

66. 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.

67. 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:

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

68. Mass is conserved in a chemical reaction.

	chromium(III) oxide	+	aluminum	$\rightarrow$	chromium	+	aluminum oxide
Mass:	34.0 g		12.1 g		23.3		?

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

69. From the Na_2X formula, X has a 2– charge. Because 36 electrons are present, X has 34 protons, 79 – 34 = 45 neutrons, and is selenium.

a. True. Nonmetals bond together using covalent bonds and are called covalent compounds.

b. False. The isotope has 34 protons.

c. False. The isotope has 45 neutrons.

d. False. The identity is selenium, Se.

70. a. Fe^{2+}: 26 protons (Fe is element 26.); protons – electrons = charge, 26 – 2 = 24 electrons; FeO is the formula because the oxide ion has a 2– charge.

b. Fe^{3+}: 26 protons; 23 electrons; Fe_2O_3

c. Ba^{2+}: 56 protons; 54 electrons; BaO

d. Cs^{+}: 55 protons; 54 electrons; Cs_2O

e. S^{2-}: 16 protons; 18 electrons; Al_2S_3

f. P^{3-}: 15 protons; 18 electrons; AlP

g. Br^{-}: 35 protons; 36 electrons; $AlBr_3$

h. N^{3-}: 7 protons; 10 electrons; AlN

71. From the XBr_2 formula, the charge on element X is 2+. Therefore, the element has 88 protons, which identifies it as radium, Ra. 230 – 88 = 142 neutrons.

72. 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.

73. The halogens have a high affinity for electrons, and one important way they react is to form anions of the type X^-. The alkali metals tend to give up electrons easily and in most of their compounds exist as M^+ cations. *Note*: These two very reactive groups are only one electron away (in the periodic table) from the least reactive family of elements, the noble gases.

74. The solid residue must have come from the flask.

75. In the case of sulfur, SO_4^{2-} is sulfate, and SO_3^{2-} is sulfite. By analogy:

SeO_4^{2-}: selenate; SeO_3^{2-}: selenite; TeO_4^{2-}: tellurate; TeO_3^{2-}: tellurite

76. The polyatomic ions and acids in this problem are not named in the text. However, they are all related to other ions and acids named in the text that contain a same group element. Because $HClO_4$ is perchloric acid, $HBrO_4$ would be perbromic acid. Because ClO_3^- is the chlorate ion, KIO_3 would be potassium iodate. Since ClO_2^- is the chlorite ion, $NaBrO_2$ would be sodium bromite. And finally, because HClO is hypochlorous acid, HIO would be hypoiodous acid.

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

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

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:

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

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

78. a. Ca^{2+} and N^{3-}: Ca_3N_2, calcium nitride

b. K^+ and O^{2-}: K_2O, potassium oxide

c. Rb^+ and F^-: RbF, rubidium fluoride

d. Mg^{2+} and S^{2-}: MgS, magnesium sulfide

e. Ba^{2+} and I^-: BaI_2, barium iodide

f. Al^{3+} and Se^{2-}: Al_2Se_3, aluminum selenide

g. Cs^+ and P^{3-}: Cs_3P, cesium phosphide

h. In^{3+} and Br^-: $InBr_3$, indium(III) bromide; In also forms In^+ ions, but you would predict In^{3+} ions from its position in the periodic table.

79. 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: $\frac{27.2\text{ g C}}{72.8\text{ g O}} = 0.374\text{ g C/g O}$

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

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

80. a. This is element 52, tellurium. Te forms stable 2– charged ions in ionic compounds (like other oxygen family members).

b. Rubidium. Rb, element 37, forms stable 1+ charged ions.

c. Argon. Ar is element 18.

d. Astatine. At is element 85.

81. Because this is a relatively small number of neutrons, the number of protons will be very close to the number of neutrons present. The heavier elements have significantly more neutrons than protons in their nuclei. Because this element forms anions, it is a nonmetal and will be a halogen because halogens form stable 1– charged ions in ionic compounds. From the halogens listed, chlorine, with an average atomic mass of 35.45, fits the data. The two isotopes are ^{35}Cl and ^{37}Cl, and the number of electrons in the 1– ion is 18. Note that because the atomic mass of chlorine listed in the periodic table is closer to 35 than 37, we can assume that ^{35}Cl is the more abundant isotope. This is discussed in Chapter 3.

ChemWork Problems

82.

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

83. 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.

84. carbon tetrabromide, CBr_4; cobalt(II) phosphate, $Co_3(PO_4)_2$;

magnesium chloride, $MgCl_2$; nickel(II) acetate, $Ni(C_2H_3O_2)_2$;

calcium nitrate, $Ca(NO_3)_2$

85. $Co(NO_2)_2$, cobalt(II) nitrite; AsF_5, arsenic pentafluoride; LiCN, lithium cyanide;

K_2SO_3, potassium sulfite; Li_3N, lithium nitride; $PbCrO_4$, lead(II) chromate

86. K will lose 1 e^- to form K^+. Cs will lose 1 e^- to form Cs^+.

Br will gain 1 e^- to form Br^-. Sulfur will gain 2 e^- to form S^{2-}.

Se will gain 2 e^- to form Se^{2-}.

87. a. False; magnesium is Mg. b. True

c. Ga is a metal and is expected to lose electrons when forming ions.

d. True

e. Titanium(IV) oxide is correct for this transition metal ionic compound.

Challenge Problems

88. 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:

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

89. 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, for example, 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).

90. For each experiment, divide the larger number by the smaller. In doing so, we get:

experiment 1	X = 1.0	Y = 10.5
experiment 2	Y = 1.4	Z = 1.0
experiment 3	X = 1.0	Y = 3.5

Our assumption about formulas dictates the rest of the solution. For example, if we assume that the formula of the compound in experiment 1 is XY and that of experiment 2 is YZ, we get relative masses of:

X = 2.0; Y = 21; Z = 15 (= 21/1.4)

and a formula of X_3Y for experiment 3 [three times as much X must be present in experiment 3 as compared to experiment 1 (10.5/3.5 = 3)].

However, if we assume the formula for experiment 2 is YZ and that of experiment 3 is XZ, then we get:

X = 2.0; Y = 7.0; Z = 5.0 (= 7.0/1.4)

and a formula of XY_3 for experiment 1.

Any answer that is consistent with your initial assumptions is correct.

The answer to part d depends on which (if any) of experiments 1 and 3 have a formula of XY in the compound. If the compound in expt. 1 has formula XY, then:

$$21\text{ g XY} \times \frac{4.2\text{ g Y}}{(4.2 + 0.4)\text{ g XY}} = 19.2\text{ g Y (and 1.8 g X)}$$

If the compound in experiment 3 has the XY formula, then:

$$21\text{ g XY} \times \frac{7.0\text{ g Y}}{(7.0 + 2.0)\text{ g XY}} = 16.3\text{ g Y (and 4.7 g X)}$$

Note that it could be that neither experiment 1 nor experiment 3 has XY as the formula. Therefore, there is no way of knowing an absolute answer here.

91. 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 combines 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.

92. 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].

93. 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 reacts, 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 × 2 = 18 atoms of H. Octane formula = C_8H_{18} and the ratio of C:H = 8:18 or 4:9.

94. 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.

Marathon Problem

95. 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:

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

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

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

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.