CHAPTER 2

ATOMS, MOLECULES, AND IONS

Questions

16. 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, etc.

17. 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.

18. 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.

19. 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 2.2, the mass ratios of oxygen that react with 1 g carbon in each compound are in a 2 : 1 ratio.

20. 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.

21. J. J. Thomson's study of cathode-ray tubes led him to postulate the existence of negatively charged particles that we now call electrons. Thomson also postulated that atoms must contain positive charge in order for the atom to be electrically neutral. 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.

22. 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.

23. 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.

24. 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.

25. 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.

26. Some properties of metals are

(1) conduct heat and electricity;

(2) malleable (can be hammered into sheets);

(3) ductile (can be pulled into wires);

(4) lustrous appearance;

(5) form cations when they form ionic compounds.

Nonmetals generally do not have these properties, and when they form ionic compounds, nonmetals always form anions.

27. Carbon is a nonmetal. Silicon and germanium are called metalloids because 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.

28. a. A molecule has no overall charge (an equal number of electrons and protons are present). Ions, on the other hand, have extra electrons added or removed to form anions (negatively charged ions) or 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; e.g., Cl^-, O^{2-}, and SO_4^{2-} are all anions. A cation is a positively charged ion, e.g., Na^+, Fe^{3+}, and NH_4^+ are all cations.

29. 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.

30. 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.

31. Statements a and b are true. Counting over in the periodic table, element 118 will be the next noble gas (a nonmetal). For statement c, hydrogen has mostly nonmetallic properties. For statement d, a family of elements is also known as a group of elements. For statement e, two items are incorrect. When a metal reacts with a nonmetal, an ionic compound is produced, and the formula of the compound would be AX_2 (alkaline earth metals form 2+ ions and halogens form 1– ions in ionic compounds). The correct statement would be: When an alkaline earth metal, A, reacts with a halogen, X, the formula of the ionic compound formed should be AX_2.

32. 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 in compounds, 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).

Exercises

Development of the Atomic Theory

33. 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. $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$. From the balanced equation, the volume of HCl produced will be twice the volume of H_2 (or Cl_2) reacted.

34. 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:

$$1\ N_2 + 3\ H_2 \rightarrow 2 \text{ product}$$

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

35. 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:

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

36. 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.

37. 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.

38. 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.

39. 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.

40. 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.

41. 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.

42. 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

43. 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.

44. 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:

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

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

The Nature of the Atom

45. From section 2.5, 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):

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

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

46. 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

47. $5.93 \times 10^{-18}\ \text{C} \times \frac{1\ \text{electron charge}}{1.602 \times 10^{-19}\ \text{C}}$ = 37 negative (electron) charges on the oil drop

48. 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.

49. sodium–Na; radium–Ra; iron–Fe; gold–Au; manganese–Mn; lead–Pb

50. fluorine–F; chlorine–Cl; bromine–Br; sulfur–S; oxygen–O; phosphorus–P

51. Sn–tin; Pt–platinum; Hg–mercury; Mg–magnesium; K–potassium; Ag–silver

52. As–arsenic; I–iodine; Xe–xenon; He–helium; C–carbon; Si–silicon

53. 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.

54. 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) has only radioactive isotopes.

55. a. transition metals b. alkaline earth metals c. alkali metals

d. noble gases e. halogens

56. 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

57. 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. ${}^{7}_{3}Li$

58. 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. ${}^{47}_{20}Ca$ f. ${}^{65}_{29}Cu$

59. 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. ${}^{16}_{8}O$

60. 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.

61. 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

62. 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

63. 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.

64. 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

65. 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: $^{118}_{50}Sn^{2+}$

66. 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: $^{32}_{16}S^{2-}$

67.

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–

68.

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–

69. 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^+.

70. See Exercise 69 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

71. a. sodium bromide b. rubidium oxide

c. calcium sulfide d. aluminum iodide

e. SrF_2 f. Al_2Se_3

g. K_3N h. Mg_3P_2

72. 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

73. a. cesium fluoride b. lithium nitride

c. silver sulfide (Silver only forms stable 1+ ions in compounds, so no Roman numerals are needed.)

d. manganese(IV) oxide e. titanium(IV) oxide f. strontium phosphide

74. a. $ZnCl_2$ (Zn only forms stable +2 ions in compounds, so no Roman numerals are needed.)

b. SnF_4 c. Ca_3N_2 d. Al_2S_3

e. Hg_2Se f. AgI (Ag only forms stable +1 ions in compounds.)

75. a. barium sulfite b. sodium nitrite

c. potassium permanganate d. potassium dichromate

76. a. $Cr(OH)_3$ b. $Mg(CN)_2$

c. $Pb(CO_3)_2$ d. $NH_4C_2H_3O_2$

77. a. dinitrogen tetroxide b. iodine trichloride

c. sulfur dioxide d. diphosphorus pentasulfide

78. a. B_2O_3 b. AsF_5
c. N_2O d. SCl_6

79. 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 tetrachloride h. sodium hypochlorite
i. barium chromate j. ammonium nitrate

80. a. acetic acid b. ammonium nitrite c. cobalt(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

Note: For the compounds named as acids, we assume these are dissolved in water.

81. 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

82. From the anion names of hypochlorite (ClO^-), chlorite (ClO_2^-), chlorate (ClO_3^-), and perchlorate (ClO_4^-), the oxyanion names for similar iodine ions would be hypoiodite (IO^-), iodite (IO_2^-), iodate (IO_3^-), and periodate (IO_4^-). The corresponding acids would be hypoiodous acid (HIO), iodous acid (HIO_2), iodic acid (HIO_3), and periodic acid (HIO_4).

83. a. SF_2 b. SF_6 c. NaH_2PO_4
d. Li_3N e. $Cr_2(CO_3)_3$ f. SnF_2
g. $NH_4C_2H_3O_2$ h. NH_4HSO_4 i. $Co(NO_3)_3$
j. Hg_2Cl_2; mercury(I) exists as Hg_2^{2+} ions. k. $KClO_3$ l. NaH

84. a. CrO_3 b. S_2Cl_2 c. NiF_2
d. K_2HPO_4 e. AlN
f. NH_3 (Nitrogen trihydride is the systematic name.) g. MnS_2
h. $Na_2Cr_2O_7$ i. $(NH_4)_2SO_3$ j. CI_4

85. a. Na_2O b. Na_2O_2 c. KCN
d. $Cu(NO_3)_2$ e. $SeBr_4$ f. HIO_2
g. PbS_2 h. CuCl
i. GaAs (We would predict the stable ions to be Ga^{3+} and As^{3-}.)
j. CdSe (Cadmium only forms 2+ charged ions in compounds.)
k. ZnS (Zinc only forms 2+ charged ions in compounds.)
l. HNO_2 m. P_2O_5

86. a. $(NH_4)_2HPO_4$ b. Hg_2S c. SiO_2
d. Na_2SO_3 e. $Al(HSO_4)_3$ f. NCl_3
g. HBr h. $HBrO_2$ i. $HBrO_4$
j. KHS k. CaI_2 l. $CsClO_4$

87. 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

88. 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. Phosphide is P^{3-}, while phosphate is PO_4^{3-}. 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.

Additional Exercises

89. 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.

90. Carbon (C); hydrogen (H); oxygen (O); nitrogen (N); phosphorus (P); sulfur (S)

For lighter elements, stable isotopes usually have equal numbers of protons and neutrons in the nucleus; these stable isotopes are usually the most abundant isotope for each element. Therefore, a predicted stable isotope for each element is ^{12}C, ^{2}H, ^{16}O, ^{14}N, ^{30}P, and ^{32}S. These are stable isotopes except for ^{30}P, which is radioactive. The most stable (and most abundant) isotope of phosphorus is ^{31}P. There are exceptions. Also, the most abundant isotope for hydrogen is ^{1}H; this has just a proton in the nucleus. ^{2}H (deuterium) is stable (not radioactive), but ^{1}H is also stable as well as most abundant.

91. $^{53}_{26}Fe^{2+}$ has 26 protons, 53 – 26 = 27 neutrons, and two fewer electrons than protons (24 electrons) in order to have a net charge of 2+.

92. 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.

93. From the Na_2X formula, X has a 2− charge. Because 36 electrons are present, X has 34 protons and 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.

94. a. Fe^{2+}: 26 protons (Fe is element 26.); protons − electrons = net charge, 26 − 2 = 24 electrons; FeO is the formula since the oxide ion has a 2− charge, and the name is iron(II) oxide.

b. Fe^{3+}: 26 protons; 23 electrons; Fe_2O_3; iron(III) oxide

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

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

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

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

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

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

95. 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)

96. 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.

97. 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.

98. 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.

99. 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 one would predict In^{3+} ions from its position in the periodic table.

100. These compounds are similar to phosphate (PO_4^{3-}) compounds. Na_3AsO_4 contains Na^+ ions and AsO_4^{3-} ions. The name would be sodium arsenate. H_3AsO_4 is analogous to phosphoric acid, H_3PO_4. H_3AsO_4 would be arsenic acid. $Mg_3(SbO_4)_2$ contains Mg^{2+} ions and SbO_4^{3-} ions, and the name would be magnesium antimonate.

101. 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

102. Mass is conserved in a chemical reaction.

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

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

ChemWork Problems

The answers to the problems 103-108 (or a variation to these problems) are found in OWL. These problems are also assignable in OWL.

Challenge Problems

109. Copper (Cu), silver (Ag), and gold (Au) make up the coinage metals.

110. Because the gases are at the same temperature and pressure, the volumes are directly proportional to the number of molecules present. Let's assume 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:

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

111. 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 × 2 = 18 atoms of H. Octane formula = C_8H_{18}, and the ratio of C : H = 8 : 18 or 4 : 9.

112. From Section 2.5 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.

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

114. The equation for the reaction would be $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$. 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 forms. 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.

115. 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).

116. 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, $\frac{1.00}{1.00 + 2y} = 0.3043$, $y = 1.14$.

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

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.

117. 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].

118. 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 experiment 1 has a formula of 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:

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

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.

Integrated Problems

119. The systematic name of Ta_2O_5 is tantalum(V) oxide. Tantalum is a transition metal and requires a Roman numeral. Sulfur is in the same group as oxygen, and its most common ion is S^{2-}. There-fore, the formula of the sulfur analogue would be Ta_2S_5.

Total number of protons in Ta_2O_5:

Ta, Z = 73, so 73 protons × 2 = 146 protons; O, Z = 8, so 8 protons × 5 = 40 protons

Total protons = 186 protons

Total number of protons in Ta_2S_5:

Ta, Z = 73, so 73 protons × 2 = 146 protons; S, Z = 16, so 16 protons × 5 = 80 protons

Total protons = 226 protons

Proton difference between Ta_2S_5 and Ta_2O_5: 226 protons – 186 protons = 40 protons

120. The cation has 51 protons and 48 electrons. The number of protons corresponds to the atomic number. Thus this is element 51, antimony. There are 3 fewer electrons than protons. Therefore, the charge on the cation is 3+. The anion has one-third the number of protons of the cation, which corresponds to 17 protons; this is element 17, chlorine. The number of electrons in this anion of chlorine is 17 + 1 = 18 electrons. The anion must have a charge of 1−.

The formula of the compound formed between Sb^{3+} and Cl^- is $SbCl_3$. The name of the compound is antimony(III) chloride. The Roman numeral is used to indicate the charge on Sb because the predicted charge is not obvious from the periodic table.

121. 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.

Marathon Problem

122. 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.