CHAPTER 2

ATOMS, MOLECULES, AND IONS

Questions

18. 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.

19. 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.

20. 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.

21. 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.

22. 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.

23. 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.

24. 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.

25. 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.

26. 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.

27. 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.

28. 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.

29. 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.

30. 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.

31. In the paste, sodium chloride will dissolve to form separate Na^+ and Cl^- ions. With the ions present and able to move about, electrical impulses will be conducted.

32. 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.

33. 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.

34. 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.

35. Statements a and b are true. Element 118, Og, is a noble gas and will presumably be 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 halo-gens 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.

36. 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

37. 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.

38. 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.

39. 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:

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

40. 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.

41. 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: $\frac{21.8\text{ g C}}{58.2\text{ g O}} = 0.375\text{ g C/g O}$

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

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.

42. Let's determine the ratios of the masses of fluorine that combine with 1 g of S:

$$\frac{1.188}{1.188} = 1.000; \qquad \frac{2.375}{1.188} = 1.999; \qquad \frac{3.563}{1.188} = 2.999$$

Because the masses of the fluorine that combine with 1.000 g of sulfur in the three compounds are all in whole number multiples, this illustrates the law of multiple proportions.

43. 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.

44. 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.

45. 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.

46. Mass is conserved in a chemical reaction.

	ethanol +	oxygen →	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

47. 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.

48. 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

49. 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):

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

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

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

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

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

50. 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

51. $5.93 \times 10^{-18}\ \text{C} \times \frac{1\ \text{electron charge}}{1.602 \times 10^{-19}\ \text{C}}$ = 37 negative (electron) charges on the oil drop

52. 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.

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

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

55. a. 6; the group 2A elements are Be, Mg, Ca, Sr, Ba, and Ra.

b. 6; the group 6A elements are O, S, Se, Te, Po, and Lv.

c. 4; the nickel family elements are Ni, Pd, Pt, amd Ds.

d. 7; the noble gas group 8A elements are He, Ne, Ar, Kr, Xe, Rn, and Uuo.

56. a. 6; the halogen group 7A elements are F, Cl, Br, I, At, and Uus.

b. 6; the alkali group 1A elements are Li, Na, K, Rb, Cs, and Fr. H is not considered an alkali metal.

c. 14; elements Ce through Lu are the 14 lanthanide series elements.

d. 40; the block of elements from Sc to Zn to Cn to Ac and back to Sc are all transition metal elements.

57. 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.

58. 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.

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

d. noble gases e. halogens

60. 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

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

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

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

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

62. 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. ${}^{65}_{29}\text{Cu}$

63. 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. ${}^{23}_{11}\text{Na}$ b. ${}^{19}_{9}\text{F}$ c. ${}^{16}_{8}\text{O}$

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

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

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

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

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

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

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

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

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

67. 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.

68. 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. ${}^{79}_{34}\mathrm{Se}$: 34 p, 45 n, 34 e
g. ${}^{79}_{34}\mathrm{Se}^{2-}$: 34 p, 45 n, 36 e

h. ${}^{63}_{28}\mathrm{Ni}$: 28 p, 35 n, 28 e
i. ${}^{59}_{28}\mathrm{Ni}^{2+}$: 28 p, 31 n, 26 e

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

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

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

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

71.

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

72.

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–

73. 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^+.

74. See Exercise 73 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

75. 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

76. 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

77. 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

78. 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.)

79. a. barium sulfite b. sodium nitrite

c. potassium permanganate d. potassium dichromate

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

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

81. a. dinitrogen tetroxide b. iodine trichloride

c. sulfur dioxide d. diphosphorus pentasulfide

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

83. 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

84. 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.

85. 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

86. 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).

87. 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

88.
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

89.
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

90.
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. $CsClO_4$

91.
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

92.
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

93. There should be no difference. The composition of insulin (the number and types of atoms) from both sources will be the same and therefore, it should have the some activity regardless of the source. As a practical note, trace contaminants in the two types of insulin may be different. These trace contaminants may be important towards the activity of insulin in the body.

94. The formula of glucose is $C_6(H_2O)_6$, or $C_6H_{12}O_6$.

95. a. $^{131}_{53}I$ has 53 protons and 131 – 53 = 78 neutrons.

b. $^{201}_{81}Tl$ has 81 protons and 201 – 81 = 120 neutrons.

96. 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.

97. $^{53}_{26}Fe^{2+}$ has 26 protons, 53 – 26 = 27 neutrons, and two fewer electrons than protons (24 electrons) in order to have a net charge of 2+.

98. 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.

99. 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.

100. 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

101. 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)

102. 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.

103. 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.

104. 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.

105. 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.

106. 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.

107. 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

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

109. 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:

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

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

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

110. a. Hydrosulfuric acid if dissolved in water or dihydrogen sulfide if a gas.

b. sulfur dioxide

c. sulfur hexafluoride

d. sodium sulfite

ChemWork Problems

111.

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

112.

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

113.

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

114. 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.

115. 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, $Ca(NO_3)_2$

116. $Co(NO_2)_2$, cobalt(II) nitrite; AsF_5, arsenic pentafluoride; LiCN, lithium cyanide;
K_2SO_3, potassium sulfite; Li_3N, lithium nitride; $PbCrO_4$, lead(II) chromate

117. 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-}.

118. 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

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

120. 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:

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

121. 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.

122. 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.

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

124. 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.

125. 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).

126. 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.

127. 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].

128. 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

129. 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-}. Therefore, 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

130. 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.

131. Number of electrons in the unknown ion:

$$2.55 \times 10^{-26}\ \text{g} \times \frac{1\,\text{kg}}{1000\ \text{g}} \times \frac{1\,\text{electron}}{9.11\times10^{-31}\ \text{kg}} = 28\ \text{electrons}$$

Number of protons in the unknown ion:

$$5.34 \times 10^{-23}\ \text{g} \times \frac{1\,\text{kg}}{1000\ \text{g}} \times \frac{1\,\text{proton}}{1.67\times10^{-27}\ \text{kg}} = 32\ \text{protons}$$

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:

$$9.35 \times 10^{-23}\ \text{g} \times \frac{1\,\text{kg}}{1000\ \text{g}} \times \frac{1\,\text{proton}}{1.67\times10^{-27}\ \text{kg}} = 56\ \text{neutrons}$$

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

132. 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.

PART II: CHAPTER DISCUSSIONS

COURSE CONTENT: DESCRIPTIVE CHEMISTRY AND CHEMICAL PRINCIPLES

There has been much discussion and an overall agreement that more descriptive chemistry needs to be included in the general chemistry course. This has been one of the major considerations in the recent curriculum reform movement. How to accomplish this goal has resulted in much less agreement. General chemistry has to move beyond a rote presentation of the facts and principles of chemistry. We feel that the facts (descriptive chemistry) and principles must be an integrated whole that is aimed at fulfilling the basic goals of the course.

The *Chemistry* textbook follows what has been for the past several years a fairly traditional order of topics. The chapters on principles (1-18) are followed by nuclear chemistry (19), a more descriptive chapter on chemistry of the elements (20), a chapter on transition metals and coordination chemistry (21), and organic chemistry and biochemistry (22). Different instructors will have different approaches to descriptive chemistry and, in fact, will probably give greatly different definitions of descriptive chemistry. We have organized the text this way because we feel this order offers instructors the greatest flexibility.

In the fifties and early sixties, general chemistry texts essentially went through the periodic table; principles were introduced as needed. Observations were made of the properties and reactions of matter; then we used chemical principles to try to systematically organize and understand those observations. This approach may closely resemble how chemists do chemistry. But is it the best way to learn chemistry? We don’t think so. We are all driven by curiosity as to how things work, and an unfortunate consequence of marching through the periodic table is to dull the students’ curiosity about why things happen, replacing it with the impression of chemistry as a vast body of unrelated facts and reactions that must be memorized. This approach tends to make learning chemistry both frustrating and uninteresting. In the sixties, a principles-dominated outline began to be used in texts, but we overcompensated. The principles were emphasized. The framework was put in place to deal with chemical facts, but all too often the facts were left out. Only the reaction $A + B \rightarrow C$ seems to have been covered. Learning chemistry was still frustrating, but in a different way from before.

We must integrate the two approaches. In truth, the principles of chemistry and descriptive chemistry are two indispensable parts of the whole. In particular, we must use all of our resources in presenting reactions. Can we honestly expect a neophyte to chemistry to appreciate the difference between

$$2H_2 + O_2 \rightarrow 2H_2O \quad \text{and} \quad A + B \rightarrow C$$

if they are only written on the blackboard or in a book? We must make use of all of the senses. Students can appreciate what $2H_2 + O_2 \rightarrow 2H_2O$ means if the instructor ignites a mixture of hydrogen and oxygen. They can see, hear, and sometimes feel what that equation represents.

We can't just talk about reactions. We must show the students reactions or, better yet, let them see the reactions on their own in the laboratory. We should choose more experiments for the labs in which students observe reactions or synthesize compounds and do fewer experiments that involve a measurement that confirms (often poorly) some principles discussed in class. The use of lecture demonstrations also allows students to see chemistry. These two approaches represent powerful tools for presenting reactions.

In the text we have tried to emphasize the framework of principles by which chemical facts can be organized. There are several important features that facilitate the integration of fact and principle.

1. To assist instructors in providing interesting and relevant classroom demonstrations of chemical phenomena for their students, the *Instructor's Annotated Edition* indicates via the marginal icon (flask with stirring rod) direct references to more than 750 demonstrations from several authoritative sources.

2. Chapter 4 presents an early, thorough discussion of solution reactions, using the classifications that chemists typically use (acid-base, precipitation, and oxidation-reduction).

3. The illustrations and color photography show the student many reactions. This merely supplements lecture demonstrations and labs but right in the text students can see the beauty of the reactions and substances composing this body of knowledge called chemistry.

4. Real examples are used to illustrate chemical principles.

5. Much of the discussion in the descriptive chapters is organized according to a framework of the early chapters. For example, Chapters 19-21 emphasize periodic relationships; the discussion of polymers in Chapter 22 emphasizes structure-property relationships.

6. Many of the end-of-chapter exercises deal with reactions or substances that are important to our lives. In the problem statement, these relationships are mentioned.

7. The end-of-chapter exercises in Chapters 19-22 are designed to illustrate the relationships between the descriptive material and the principles chapters. Energy relationships (particularly bond energies), periodicity, structure, and

equilibrium are emphasized.

In the following chapter discussions, we will further point out how descriptive material from the later chapters can be brought into discussion of principles. The organization of the text allows each instructor greater flexibility in presenting the material. The unique features of the text allow for a more thorough integration of principles and facts. Chemistry isn't just principles; it isn't just reactions and properties of elements and compounds. It is an amalgam of both. In the text we have tried to bring both together, while allowing instructors maximum flexibility.

CHAPTER ONE: CHEMICAL FOUNDATIONS

Chapter Learning Goals:

Section One:	To appreciate the importance of creative problem solving.
Section Two:	To identify the principal operations and limitations of the scientific method.
Section Three:	To describe the SI system of units and prefixes.
Section Four:	To identify causes of uncertainty in measurement. To show how significant figures are used. To compare precision and accuracy in measurement.
Section Five:	To show how to determine the number of significant figures in a calculated result.
Section Six:	To show a general method of solving problems.
Section Seven:	To show how to convert units between the English and metric systems.
Section Eight:	To demonstrate conversions among the Fahrenheit, Celsius, and Kelvin temperature scales.
Section Nine:	To illustrate calculations involving density.
Section Ten:	To show how matter can be classified into subgroups.

The time spent on the first two chapters may show the greatest differences from class to class. Much of the first chapter is probably review for students with good high school math and chemistry backgrounds.

This chapter lays the foundation for dealing with measured quantities and performing calculations. Instructors wishing to treat uncertainty in greater detail can discuss the section Uncertainties in Measurements, from Appendix One. This section of the appendix could also be used in the laboratory.

Instructors should be careful to point out that the Sample Exercises are worked in a different manner from what students should do. Intermediate answers are rounded off to show the correct number of significant figures at each stage of the calculation. These rounded values are then used to complete the calculation. Students, when working problems, should round only at the end. In the *Solutions Guide*, we have followed the same convention as the text and have rounded at the points where intermediate answers are shown. If this would result in excessive round-off error, we have carried extra digits and noted this in the solution. Most commonly we have carried extra digits when solving two simultaneous equations and in doing equilibrium problems.

One of the primary reasons for the discussion of units is to introduce the student to the use of dimensional analysis as a problem-solving technique. In our classes we emphasize using units as a check. Some students may tend to be sloppy about units in calculations because they are familiar with the quantities being used. The nice feature of dimensional analysis is that it works well even if we don't have a good intuitive feel for the quantities and units encountered in a particular problem. On the other hand, it is important to remember that correct use of dimensional analysis does not mean that a student has a conceptual understanding of the chemical concepts in a particular problem. One of the reasons we have included the Active Learning Questions is so the students have a chance to vocalize their ideas about the concepts covered, and we have a better chance to try to understand these.

Teaching Tips

- Connecting the real world to the atomic/molecular world is difficult for many students. Encourage students to visualize the atomic/molecular world by drawing pictures. The student responses will give you an idea of their entering perceptions. Don't worry about some of the details such as subatomic particles here. For example, point out the representations of hydrogen, oxygen, and water in Section 1.1.

- Consider spending some time discussing Figure 1.4. Some students believe that theories become laws once a theory has been accepted, but this figure clearly shows laws and theories as two separate entities. Theories never become laws. Laws tell us what happens, and theories are our attempts to explain why. In terms of a "scientific method", students should understand that much of scientific thinking is logical or analytical thinking.

- To illustrate uncertainty in measurements bring to class several pieces of glassware used to measure volume. Show the students the glassware. Draw the shape of a representative piece of the glassware on the board marking the drawing to show the meniscus of a liquid between two of the lines. Since the volume falls between the lines there is uncertainty involved in the reading. Have the students write down the result of the measurement. Then compare results. The number recorded for this measurement will vary from student to student only in the last digit. Thus there is only one uncertain digit.

- A relatively easy way to determine the number of significant figures in a number is to write the number in scientific notation. In this notation there are no leading zeros, all zeros in the middle of the number are significant, and trailing zeros are only recorded if they are significant.

- To illustrate the difference between a physical and chemical change compare representation (b) and (c) from Figure 1.11 (a physical change from liquid water to gaseous water) to the chemical change when liquid water is electrolyzed to produce gases H_2 and O_2. Use the equation in the text to represent the electrolysis of water chemical change.

CHAPTER TWO: ATOMS, MOLECULES, AND IONS

Chapter Learning Goals:

Section One:	To give a brief account of early chemical discoveries.
Section Two:	To describe and illustrate the laws of conservation of mass, definite proportion, and multiple proportions.
Section Three:	To describe Dalton's theory of atoms and show the significance of Gay-Lussac's experiments.
Section Four:	To summarize the experiments that characterized the structure of the atom.
Section Five:	To describe features of subatomic particles.
Section Six:	To introduce basic ideas of bonding in molecules. To show various ways of representing molecules.
Section Seven:	To introduce various features of the periodic table.
Section Eight:	To demonstrate how to name compounds given their formulas and to write formulas given their names.

Chapter 2 is the second of the two introductory chapters. As with Chapter 1, the length of time spent on this chapter will depend on the background of the students. It goes over the background in chemical topics that provide the foundation for the rest of the course.

This chapter takes a historical approach to the development of chemistry. It begins with a discussion of the discoveries leading to Dalton’s atomic theory, continues with the experiments elucidating the structure of the atom, and ends with chemical nomenclature and an introduction to the periodic table.

The section on nomenclature and formula writing in the laboratory manual by Hall can be effectively used at this point. Instructors may also wish to discuss nuclear decay in more detail at this time. Chapter 19, Nuclear Chemistry, can be covered here if instructors prefer. However, if the chapter is presented early, the sections on decay kinetics and thermodynamic stability of the nucleus would have to be delayed and could be covered with Chemical Kinetics (Chapter 12) or Spontaneity, Entropy, and Free Energy (Chapter 17).

Teaching Tips

- Consider using an analogy between the letters in the alphabet and words and the atoms of the various elements and compounds. In the same way we can make many words from 26 letters, millions of compounds can be made using only the 100 or so elements. Putting the same elements together in different ways results in different compounds with very different properties. For example, the words “ADD” and “DAD” each consist of the same letters. However, the order of the letters gives them different meanings. This discussion will help later (for example, in Chapter 3 you can use this analogy to help the students understand why we do not change subscripts in a chemical formula).

- Students often have difficulty understanding what is meant by chemical terms such as element. It is important to emphasize the various ways we use the term element in chemistry. When we say element we might mean a single atom, we might mean a molecule such as N or O or we might mean a large sample such as a bar of aluminum or the graphite (“lead”) in a pencil. Be sure to help students realize that chemists use terms in many ways and they need to look at the context of the word to be sure that they understand what is intended. Chemists are accustomed to thinking about things at the macroscopic and microscopic level simultaneously. Encourage students to think about what they read and to consider how the terms are being used.

- Sections 2.2 and 2.3 provide another opportunity to discuss the difference between laws and theories. Dalton’s atomic theory is relatively simple in scope, and this is as it should be. John Dalton was trying to explain laws such as the law of constant composition. The success of this makes the model successful, but not absolutely correct (a model is always a simplification). For example, Dalton’s theory does not explain questions such as “Why/how do atoms stick together to form molecules?”, and “Why/how do molecules stick together to form liquids and solids?”. But no model answers all questions.

- The discovery of the proton leads to an excellent example of how and why models change. Dalton’s model of the atom did not account for isotopes since he assumed that all atoms of an element are exactly alike. As knowledge expands models change to accommodate the new information

- Students should learn the names of the common groups on the periodic table (see Figure 2.19). This will simplify your discussion later in the course. Reasons behind the structure of the periodic table and trends it shows are not introduced until Chapter 7. Focus the students’ attention of the location of groups of elements and the separation between metals and nonmetals.

- Emphasize that ions are formed by the gain and loss of electrons. This is difficult for students because they tend to think about gain and loss using positive numbers. It is very important that they understand from the beginning that a positive ion is formed by losing electrons and a negative ion is formed by gaining electrons. In chemistry the positive number (protons) remain the same.

- Be sure to point out to students that atoms do not spontaneously gain or lose electrons. The gain and loss of electrons are always paired. One atom loses electrons simultaneously with another atom gaining electrons.

- Stress that learning to name compounds does not consist of memorizing a seemingly endless list of chemicals. There are systematic rules for naming compounds and, by knowing only a few rules, the students can name most any compound they will encounter in this course.

- The students should also understand that we keep the rules as simple as possible (much the same way we keep our scientific models as simple as possible). We only add complications (such as prefixes and Roman numerals) when it is required for clarity. For example, the name sodium(I) chloride is not necessarily wrong, merely redundant. Therefore, the name sodium chloride suffices.

- The students can determine the charges of all the ions in Table 2.3 from their positions on the periodic table, with the exception of Ag^+. The students should appreciate that the periodic table contains a great deal of information. It is a valuable resource.

- Make a clear distinction between Type I and Type II cations. This reinforces the reasons for the Roman numerals in the names of the compounds. Ask students, for example “How is ‘iron oxide’ different from ‘magnesium oxide’?”

- Students often make the mistake that the Roman numeral tells us that number of ions present in the compound. This is because in many cases the Roman numeral is the same as one of the subscripts. Use an example such as iron(II) oxide with the formula FeO to show that the Roman numeral relates only to the charge.

- For Type III compounds, consider starting with examples such as “NO” and “NO_2”. Ask the students to name these compounds. This will lead naturally to a discussion of prefixes. Like Roman numerals, prefixes are used only for clarity. For example, there are two possible “carbon oxides” and we must use prefixes to differentiate between these (carbon monoxide and carbon dioxide). We only make the rules more complicated when it is required.

- Figure 2.22 and Figure 2.23 are convenient flow charts to help students systematically name compounds. Students need to be encouraged to use this type of device to help them think through a difficult problem.

- Students should think of polyatomic ions as a “unit”. Many students know that if NaCl is dissolved in water, Na^+ ions and Cl^- ions are present. However, students are often confused about the ions present when $NaNO_3$ is dissolved. Instead of the NO_3^- ion students often think individual nitrogen and oxygen ions are present.

CHAPTER THREE: STOICHIOMETRY

Chapter Learning Goals:

Section One: To describe the modern atomic mass scale and explain how atomic masses are determined experimentally.
Section Two: To explain atomic mass and its experimental determination.
Section Three: To explain the importance of the mole concept.
To show how to convert among moles, mass, and number of particles for a given sample.
Section Four: To show how to calculate values for molar mass.
To show how to convert among molar mass, moles, and number of particles in a given sample.
Section Five: To describe a conceptual problem solving approach to chemistry.
Section Six: To demonstrate the calculation of the mass percent of a given element in a compound.
Section Seven: To demonstrate the calculation of the empirical formula of a compound.
To show how to obtain the molecular formula, given the empirical formula and the molar mass.
Section Eight: To identify the characteristics of a chemical reaction and the information given by a chemical equation.
Section Nine: To show how to write a balanced equation to describe a chemical reaction.
Section Ten: To show how to calculate the masses of reactants and products using the chemical equation.
Section Eleven: To show how to recognize the limiting reactant.
To demonstrate the use of the limiting reactant in stoichiometric calculations.

Chapter 3 deals with the fundamental measurement unit in chemistry, the chemical mole. The law of conservation of mass is the unifying principle of the chapter. The chapter deals with compounds first and then reactions.

An important point that can be made in lecture is the convenience of the mole as a unit. Chemical reactions, chemical formulas, and structures of molecules focus on numbers of atoms, molecules, or ions. Atoms are so small that we cannot see them to count them. Even if we could see them, the numbers that we would encounter are so large it would take an immense period of time to count them. The mole provides a unit that allows us to connect the number of atoms (what we are interested in knowing) to something we can measure (such as mass). This approach makes it easy to introduce molarity in Chapter 4 as a similar type of unit. Molarity enables us to convert between the quantity we need to deal with in reactions (numbers of moles) and the most convenient measurement we can make for the amount of a solution (volume).

Teaching Tips

- Section 3.5 discussing a conceptual problem solving strategy for the students. It offers suggestions on thinking about the problem so that they learn how to solve problems in general, not simply use the right algorithm for a given problem. It is well worthwhile to spend some time with the students discussing this.

- The concept of counting atoms from the mass of a sample can be difficult for students to understand. Develop the idea slowly. Many students learn to work the problems without really understanding this section. This leads to trouble later when they try to solve more complex problems. Return to the candy analogy to assist in developing this concept. You can also use a hardware store analogy of buying nails by weight, or a banking analogy of counting coins into rolls by weighing them.

- The idea that different samples with identical mass ratios contain the same number of objects is a difficult one for the students. Consider using the following example.

  Suppose we have two blocks, a red block and a yellow block. The red block weighs 1.0 ounce, and the yellow block weighs 4.0 ounces. Now suppose we have 16 of each color block. What is the mass of each sample? The sample of red blocks weighs 16.0 ounces and the sample of yellow blocks weighs 64.0 ounces. But note that 16.0 ounces is also 1.0 pound, thus 64.0 ounces is 4.0 pounds. The relative masses of the blocks are:

	one block	sixteen blocks
red	1.0 ounce	1.0 pound
yellow	4.0 ounces	4.0 pounds

  The relative masses stay the same (1:4) but the units are changed.

  You can make an analogy between ounces and pounds in this case and amu's and grams in the case of the periodic table. In the blocks example, the number 16 is analogous to Avogadro's number.

- It is important for students to realize that a mole really describes the number of objects present, just as a dozen means 12. This connection between a mole as a number and the mass of a sample is essential because we count atoms by weighing samples containing large numbers of them.

- Notice that the text defines the mole differently from the SI definition: “the mole is the amount of substance that contains as many entities as there are in exactly 0.012 kg of carbon 12”. The text definition is “the mole is the number equal to the number of carbon atoms in 12.01 grams of carbon”. We believe the text definition is easier for the students to understand at this point as it more strongly emphasizes both that the mole is fundamentally a number and that the periodic table contains average atomic masses.

- In Section 3.9 students learn that a chemical reaction involves rearrangement of the elements. It is very helpful to use simple models of the reactants that you can take apart and rearrange to show students that atoms are conserved in a chemical reaction (toothpicks and gumdrops or clay work well if you do not have ball-and-stick models). Use the molecular level graphics of chemical reactions in Section 3.9 to aid your discussion.

- The text lists three steps for writing and balancing chemical equations. Step 3 tells the students to start “with the most complicated molecule(s).” For the reaction

  $$C_2H_5OH(g) + O_2(g) \rightarrow CO_2 + H_2O(l)$$

  consider balancing the equation in the following way:

  $C_2H_5OH(g) + O_2(g) \rightarrow 2CO_2 + H_2O(l)$ (carbon is balanced)

  $C_2H_5OH(g) + 2O_2(g) \rightarrow 2CO_2 + H_2O(l)$ (oxygen is balanced)

  $C_2H_5OH(g) + 2O_2(g) \rightarrow 2CO_2 + 3H_2O(l)$ (hydrogen is balanced, oxygen now unbalanced)

  $C_2H_5OH(g) + 3O_2(g) \rightarrow 2CO_2 + 3H_2O(l)$ (the equation is balanced)

  Notice that we eventually get the right answer (so trial and error is a reasonable method). However, it would have been better not to have balanced the oxygen as the second step. This is why Step 3 tells the students to start with the most complicated molecule. Encourage the students to leave the elements and diatomic molecules until last when balancing equations. This simplifies the trial and error process.

- Emphasize that the coefficients in a balanced equation represent the mole *ratios*. An individual coefficient is meaningless (just as an amount of a single ingredient in a recipe is meaningless without the other amounts).

- It is useful to use a real life example when introducing limiting reactants. Students seem to relate easily to kitchen experiences yet have difficulty applying the same concepts to chemical equations. By working with real life experiences first and addressing the concepts in this context students have an opportunity to become comfortable with the concept before applying it to an unfamiliar setting.

- Using a concrete analogy (such as making sandwiches) helps the students better understand the calculations. This is especially true for problems involving a limiting reactant. Problems in which a reactant is limiting are conceptually the same as the cases in which no reactant is limiting. However, students seem to have much more difficulty dealing with limiting reactants. For example, some students calculate the amount of product that could be made from each reactant and then add them. Or, they add the moles of each given reactant for the total. They almost never make these types of mistakes in the analogies, though, so it is helpful to be very explicit about these points

- Use molecular diagrams like those shown in Section 3.11 of the text to make counting molecules more concrete (and similar to real life experience).

CHAPTER FOUR: TYPES OF CHEMICAL REACTIONS AND SOLUTION STOICHIOMETRY

Chapter Learning Goals:

Section One:	To show why the polar nature of water makes it an effective solvent.
Section Two:	To characterize strong electrolytes, weak electrolytes, and non-electrolytes.
Section Three:	To define molarity and demonstrate calculations involving the composition of solutions.
Section Four:	To introduce several types of solution reactions.
Section Five:	To show how to predict whether a solid will form in a solution reaction.
Section Six:	To describe reactions in solution by molecular, complete ionic, and net ionic equations.
Section Seven:	To demonstrate stoichiometric calculations involving precipitation reactions.

Section Eight: To show how to perform calculations involved in acid-base volumetric analysis.
Section Nine: To characterize oxidation-reduction reactions.
To describe how to assign oxidation states.
To identify oxidizing and reducing agents.
Section Ten: To describe the oxidation states method for balancing oxidation reduction reactions.

The placement of this chapter differs from its placement in many general chemistry texts. We feel there are several good reasons for including these topics at this point and many other texts have come to agree with us.

A thorough discussion of the types of reactions in solution allows for an early introduction of descriptive chemistry. Descriptive chemistry is very important in the first-year course. Although virtually every instructor will give a different definition of descriptive chemistry and a different list of topics to be covered, a thorough discussion of chemical reactions in aqueous solution is probably central to all of those definitions and lists. Hence, the expanded discussion of reactions at an early point.

Several years ago texts began introducing molarity and solution stoichiometry in the stoichiometry chapter to give flexibility to the lab program. The expanded discussion in Chapter 4 gives a great deal of flexibility to the lab. The table on scheduling lectures and labs in Part I amply illustrates this point. The stoichiometry for all of the common types of solution reactions is discussed. Students can begin to see all types of reactions in the lab very early in the course. The expanded discussion in the text will help the student focus on what is actually going on in a solution, what species are really present, and how they interact with one another.

A chemical equation written on the page of a textbook probably has little meaning to a neophyte. Students must have the opportunity to see reactions. Lecture demonstrations and the color photographs in the text help, but the most useful place for seeing chemistry is by doing chemistry in the laboratory. Chapter 4 expands the vista of what can be done in the lab.

In Chapter 4, reactions are classified as precipitation, acid-base, and oxidation-reduction. The use of terms such as metathesis, combination, displacement, double displacement, and so forth, is avoided because these terms are not used by chemists. These terms can be confusing. For example, are

$$\mathrm{AgNO_3(aq) + HCl(aq) \rightarrow AgCl(s) + HNO_3(aq)}$$

and

$$\mathrm{Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)}$$

the same or different? Is the first a double displacement and the second a combination? We strongly feel reactions should be classified on the basis of what can be observed in the lab and what is happening chemically, and not on how an equation is written on paper.

We think that an early and thorough discussion of the topics in Chapter 4 will lay a strong foundation for integrating descriptive chemistry with chemical principles throughout the rest of the course. However, instructors may wish to postpone some of the topics until later. The sections on molarity and solutions in general can be used as part of a discussion of stoichiometry or postponed until Chapter 11 is covered. Acid-base and precipitation reactions can be delayed until Chapters 14, 15 and 16. The oxidation-reduction section can be covered with Chapter 18, Electrochemistry. The sections are quite independent and delay will not cause disruption of continuity.

At the beginning of the end-of-chapter exercises in Chapters 11 and 18, there are some problems that review the topics in Chapter 4. Thus, if you decide to do all of Chapter 4 early, the students can review the material later by doing these problems. We recommend doing the entire chapter early if at all possible. We think you will find it quite effective as a complement to the lab, which really must become a more important component of the general chemistry course.

Teaching Tips

- Molarity, like density (and molar mass), is a ratio of two numbers. Students often believe that if, for example, solution A has a greater concentration that solution B, then solution A must have a greater number of moles of solute than solution B. This is not necessarily the case. For example, present the following situation:

  Solution A: dissolve 5 moles of sugar in 10 L of solution: concentration = 0.5 M

  Solution B: dissolve 1 mole of sugar in 1 L of solution: concentration = 1.0 M

  Solution B has fewer moles of sugar, but a greater concentration than solution A.

- The key concept in dilution is that the concentrated solution has the same number of moles as the dilute solution. Students are confused about this because they think of dilute as containing less solute than concentrated. In this case the solution has increased in volume and the number of moles of solute have remained the same. Stress that concentration is the amount of solute per volume available.

- The formula $M_1V_1 = M_2V_2$ works well but some students memorize it without understanding why it works. Emphasize what each of the variable means, and that it only works because the number of moles of solute is the same before and after dilution. Consider asking a question in which the students are asked to solve for the volume of water that must be added to achieve a dilution. In this way, the students need to think about what they are solving for instead of just plugging numbers into an equation.

- Students often conserve subscripts in precipitation reactions. Emphasizing what the solutions look like at a molecular level, and that the products must be electrically neutral, helps the students understand why subscripts need not be conserved.

- Discuss the importance of oxidation-reduction reactions in daily life. Batteries, corrosion, and combustion are all oxidation-reduction reactions.

- Point out to the students that the loss or gain of electrons can only take place if there is another element present to accept or donate electrons.

- Some students wonder why the term “reduction” is used to describe a *gain* of electrons. When you discuss oxidation states in Section 4.9, emphasize that when an atom is reduced its oxidation state is reduced.

- It is easiest to introduce the oxidation states for ions before discussing covalent molecules. Since the charge on the ion is the same as the oxidation state this makes sense to the students. Emphasize that the charge on an ion is acquired form a transfer of electrons. This establishes the idea of assigning electrons to a particular atom. Then move to covalent molecules and begin assigning oxidation states.

- Students do need to learn the rules for assigning oxidation states. It is a good idea to give them a number of practice problems in assigning oxidation states to all of the elements in a compound.

CHAPTER FIVE: GASES

Chapter Learning Goals:

Section One:	To demonstrate atmospheric pressure and explain how barometers work. To define the various units of pressure.
Section Two:	To describe certain laws that relate the volume, pressure, and temperature of a gas and to do calculations involving these laws.
Section Three:	To define the ideal gas law. To show how to do calculations involving the ideal gas law.
Section Four:	To define the molar volume for an ideal gas. To define STP. To show how to do stoichiometric calculations for reactions involving gases. To show how to calculate molar mass from gas density.
Section Five:	To state the relationship between partial pressures and total

	pressure and between partial pressure and mole fraction.
	To show how to obtain the molecular formula, given the empirical formula and the molar mass.
Section Six:	To present the basic postulates of the kinetic molecular theory.
	To define temperature.
	To show how to calculate and use root mean square velocity.
Section Seven:	To describe effusion and diffusion.
	To show the relationship between effusion and diffusion.
Section Eight:	To describe how real gases deviate from ideal behavior.
	To show how van der Waals's equation allows for real conditions.
Section Nine:	To characterize several real gases.
Section Ten:	To characterize the composition of the atmosphere.
	To describe some of the chemistry of air pollution.

The placement of this chapter is consistent with an historical approach. The studies of gases were fundamental in the development of the modern atomic theory. Boyle's law is consistent with an atomistic view of matter and not the Aristotelian view of matter. This material also provides an excellent example of how models are developed and used, an emphasis throughout the text. Consistent with the emphasis of other chapters, we have increased the pictorial representations of gases and their reactions.

This chapter discusses the bulk of behavior of gases (gas laws and their applications) and follows with a microscopic model (the kinetic molecular theory) that describes this bulk behavior. The kinetic molecular theory is discussed in a more quantitative manner in Appendix Two. The behavior of real gases can lead logically to a discussion of atomic and molecular structure. One of the causes for deviation from ideal behavior is the interactions arising from interparticle attractions. To understand these attractions we must consider the detailed electronic structure of atoms and molecules rather than treat them as small, billiard ball-like particles.

The other logical place to deal with gases is to cover them between Chapters 9 and 10. After we discuss the microscopic structure of matter, we deal with the bulk physical properties of matters. Gases, liquids, and solids can be taken together at this point.

Teaching Tips

- This chapter emphasizes the scientific method in action. Focus on the way that observations lead to natural laws, and how we try to explain these laws with theories or models. This is a good time to review the difference between laws and models. The gas laws allow students to predict the behavior of gases, while the kinetic molecular theory explains why gases behave in the manner they do.

- It is important for the students to understand that while the laws are based on observations the laws do not explain why the observation occurs. Laws can help us make predictions, but they are not explanations. Knowing the law of gravity, for example, allows us to predict that if we drop a pen it will fall to the ground. It does not explain why the pen fell and does not explain to us what gravity is, or how gravity “works”. To explain the observations, we develop models or theories. We want the model to be as simple as it can be in order to answer the questions we wish to answer.

- You cannot overemphasize that for Boyle’s law to hold true the temperature and amount of gas must be held constant. Certain activities, such as filling a tire or a basketball with air, seem on the surface to be at odds with Boyle’s law. Students often reason this way: “If I add more gas to a container such as a basketball, the volume and pressure both increase. Thus, as pressure increases, so does volume.” Because of this, students may believe that pressure and volume are directly related.

- Students sometimes confuse the idea of an inverse relationship with that of a line having a negative slope. A linear graph with a negative slope is a direct relationship between the variables. To make this distinction clear for these students focus on the idea that inverse means the relationship varies as the reciprocal (1/x). Have the students focus on the shape of the graph in Figure 5.5a. Ask the students to tell you what will happen to the pressure as the volume becomes extremely large (the pressure approaches zero, but will never reach zero). As the volume gets smaller, the pressure increases (as the volume goes to zero, the pressure goes to infinity). Thus, the graph must not touch the axes – it is a hyperbola not a straight line.

- At this point all four variables (P, V, T, and n) have been used in gas law problems. This is an excellent time to focus on problem solving skills. Be sure to have the students write down what they know from the problem, what they are looking for and what hey need to know to find it.

- Encourage the students to use the Ideal Gas Law to solve most gas law problems, even when some conditions are not changing. All the other gas laws can be derived from the Ideal Gas Law. If students understand how to do this they need not memorize all the other gas laws. Encourage the use of finesse and understanding to replace the brute force method of memorizing formulas and plugging numbers into them.

- When introducing Dalton’s Law of Partial Pressures encourage students to think about the mixtures of gases at the molecular level. The presence of additional gases does not affect the number of times a particular gas molecule hits the side of the container (since it is mostly empty space anyway). This helps the students understand why the pressures are additive. The observations summarized by Dalton's law of partial pressures are extremely important when developing a theory to explain the gas laws and we will develop a theory of gases that is dependent on these two statements.

- The postulates for the kinetic molecular theory form the basis for understanding the behavior of gases. Encourage students to develop a mental movie of an ideal gas on a molecular level that can form the basis for understanding the theory.

- It is also advisable to stress that we are talking about ideal gases. For example, we know real gas particles exert forces on one another as evidenced by steam condensing on a cool mirror during a hot shower. However, it is important for students to understand that although our model for gases assumes ideal behavior and gases are not ideal we can use the model to predict behavior and understand gases. The difficulties with real gases surface in small volumes and at large pressures. The ideal gas law still allows us to predict the behavior of gases under most everyday conditions.

- Once students see the molecular level action of gases with changing temperature proceed to investigate and develop models for the molecular level to illustrate the P, T and V, T relationships. This is an opportune time for the students to begin to think as a chemist. Observing gases at the macroscopic level allows prediction of behavior while developing a model at the microscopic level allows us to understand why gases behave as they do.

- Consider discussing a semi-quantitative relationship between pressure and volume for mathematically inclined students. Imagine a sample of gas in a cubic container with sides of 1 ft. If we transfer this sample to a cube with sides of 2 ft, the gas particles will have to travel (on average) twice as far from one wall to the next, and there will be half as many collisions (decreasing the pressure by a factor of 2). In addition, the area of each wall will be four times as large (4 ft^2 vs. 1 ft^2). Because pressure is a measure of force per area, this will cause the pressure to be decreased by a factor of 4. Thus, the pressure should be reduced overall by a factor of 8 (2 x 4). This is consistent with Boyle's law since the second container has a volume 8 times larger than the first container.

CHAPTER SIX: THERMOCHEMISTRY

Chapter Learning Goals:

Section One:	To describe the energy flow between a system and its surroundings. To discuss the first law of thermodynamics. To show how to calculate the work that results from changing the volume of a gas at constant pressure.

Section Two:	To define enthalpy and demonstrate calculations of the change in enthalpy in a chemical reaction. To show how a change in enthalpy is measured by calorimetry.
Section Three:	To discuss the characteristics of enthalpy changes. To show how to calculate ΔH for a chemical reaction.
Section Four:	To define standard states. To show how to use standard enthalpies of formation to calculate ΔH° for a reaction.
Section Five:	To discuss fossil fuels and the effects of their use on climate.
Section Six:	To discuss energy alternatives. To compare the available energy of various fuels.

Once again, there are other logical alternatives to the placement of this chapter. Many instructors like to treat all of thermodynamics at once. Thus this chapter can be combined with Chapter 17 and covered either here or in the second semester of the one-year course.

We have tried to strike a compromise. Students usually have an intuitive feel for energy. Putting that intuition into a formal discussion is very helpful. In the lab the students will often encounter reactions involving the generation or absorption of heat. Energy is often involved in the processes discussed in connection with atoms and bonding (e.g., ionization energy, electron affinity, bond energy). In addition, a discussion of energy allows inclusion of some interesting and important descriptive chemistry, such as that given in the chapter on the energy crisis and alternative fuels. Some discussion of energy is useful in the first part of the year. We have included the first law (conservation of energy) and enthalpy in Chapter 6. We can then talk about energies, thermodynamic cycles (such as the Born-Haber cycle), and bond energies in Chapters 7 and 8. Topics that students find more difficult conceptually, such as entropy and free energy, are deferred until later in the text.

We recommend two things for instructors who choose to cover all of thermodynamics later in the course. First, look at heat as a stoichiometric quantity and present a general discussion of energy (Sections 6.1 and 6.2). Energy concepts are used heavily in Chapters 7, 8, 10, and 11; a brief introduction to energy would make these later discussions more clear. Also, consider saving the detailed discussion of bond energies (Section 8.8) for the later, comprehensive coverage of thermodynamics.

Teaching Tips

- Another example of a state function is displacement. For example, if two students go from Chicago to New York, they are both the same distance away from where they started initially. However, they may have taken different routes; one student may have traveled a greater distance, but they end up with the same displacement.

- Students have difficulty making a distinction between temperature and heat. It is important to talk about heat as a flow of energy and temperature as a measure of the “vigor” of the random motions of the components of a substance. Thermal energy is the energy a substance contains due to the random motions of its components. Heat is the way that thermal energy is transferred from one object to another.

- Thermodynamic quantities have two parts: the magnitude and the sign. It will be a new idea for many of the students to think of the sign as a representation of direction. This idea is a key concept on which to focus attention. Emphasize that the sign convention is arbitrary. That is, since chemists chose to take the system's point of view, all signs are determined relative to the system.

- Enthalpy has no physical meaning; it is a mathematic definition. At constant pressure, enthalpy shares the same meaning as heat.

- Once students understand that energy changes can be measured in a calorimeter, the next logical question is: “How can anyone measure all of the energy changes in the world? What if the reaction is unpleasant or results in an explosion?” To answer this question, Hess’s law is useful. It provides an excellent example of how a chemist can extend knowledge by applying information gathered from simple systems to more complex systems.

- An analogy you can use with Hess's law is paying for something at a store. If the net result is to pay the store $8, this can be achieved in many ways. For example, you can give the store one $5 bill and three $1 bills; or you can pay with a $10 and get $2 in change. No matter how many steps are involved in the process, the payment is the same.

CHAPTER SEVEN: ATOMIC STRUCTURE AND PERIODICITY

Chapter Learning Goals:

Section One:	To characterize electromagnetic radiation in terms of wavelength, frequency, and speed.
Section Two:	To introduce the concept of quantized energy. To show that light has both wave and particulate properties. To describe how diffraction experiments were used to demonstrate the dual nature of all matter.
Section Three:	To show that the line spectrum of hydrogen demonstrates the quantized nature of the energy of its electron.
Section Four:	To describe the development of the Bohr model for the hydrogen atom.

Section Five:	To show how standing waves can be used to describe electrons in atoms. To describe the Heisenberg uncertainty principle. To explain the significance of electron probability distributions.
Section Six:	To explain the quantum numbers n, l, and m_l.
Section Seven:	To describe the shapes of orbitals designated by s, p, d, and f and to discuss orbital energies.
Section Eight:	To define electron spin and the electron spin quantum number. To explain the Pauli exclusion principle.
Section Nine:	To show how the quantum mechanical model can be applied to atoms besides hydrogen.
Section Ten:	To trace the development of the periodic table.
Section Eleven:	To explain the Aufbau principle.
Section Twelve:	To show general trends in ionization energy, electron affinity, and atomic radius in the periodic table.
Section Thirteen:	To show what types of information can be obtained from the periodic table.

This chapter combines into one what many texts do in two. Particular attention is paid to the relationship between models and facts. The order of topics is very similar to texts that treat electronic structure and the periodic table separately. The nature of light and the quantum mechanical model of the hydrogen atom are discussed. Then the historical development of the periodic table is followed by a discussion of the quantum mechanical model of polyelectronic atoms and how the model fits the periodic table. In discussing polyelectronic atoms, observation and the periodic table are emphasized. The concept of shielding is used to rationalize the periodic trends in size, ionization energy, and electron affinity. At first this may seem to be too sophisticated and abstract an approach for first-year chemistry students. However, the approach is very pictorial and easy to comprehend. We have found this approach to work well in our classes. Emphasis can be shifted at the instructor's option between models and the experimental facts. Combining the atomic theory with the periodic table makes this easier.

The last part of the chapter discusses the periodic trends and properties of the alkali metals. This presents the instructor with another opportunity to integrate more descriptive chemistry in with the chemical principles. Chapter 20 is organized to discuss the elements by periodic group. Instructors may choose other groups from these chapters to illustrate chemical periodicity.

A further point should be made concerning atomic radii. The volumes given in Chapter 7 are for covalent radii. The covalent radii of He, Ne, and Ar are estimates. All comparisons of atomic sizes to each other can be done consistently with this set of radii.

In Chapter 8, instructors may want to compare the sizes of a series of ions to the isoelectronic noble gases. The radii of the noble gases for this comparison should be the univalent crystal radii. These values are: He = 93 pm, Ne = 112 pm, Ar = 154 pm, Kr = 169 pm, and Xe = 190 pm. Huheey[1] and Pauling[2] provide detailed information on uses of ionic radii and tables of their values.

Teaching Tips

- Students often consider the forms of electromagnetic radiation outside the visible range (radio, X-rays, infrared, microwaves) to be entirely different from light. Use Figure 7.2 (the electromagnetic spectrum) to show the range of frequencies for each type of radiation.

- A commonly used analogy for the energy levels within an atom is the staircase. A person can move from one step to another or even move up 2 or 3 steps and down a similar number. It is not possible however to move up or down a part of a step. Within atoms electron energies are quantized – the energy levels are like stairsteps. Electrons can change only between established energy levels, not in between. Compare this to using a ramp which is like continuous energies.

- Be sure to discuss how hydrogen can have multiple lines even though it has only one electron. Students often forget that they are observing multiple atoms simultaneously when observing a line emission spectrum.

- Point out that the quantized nature of atoms is a surprise and non-intuitive. It is comforting to students to realize that this topic is complicated even to people who have been studying it for awhile.

- This is an excellent time to discuss the colors of streetlights. The mercury vapor lights have a blue cast, while the sodium lights are yellow and neon lights are more orange. You may also want to discuss the difference between an incandescent light and a vapor light. Many students think all lights are the same. They may not realize that an incandescent lamp produces a continuous spectrum since the tiny filament is radiating all wavelengths of visible light while vapor lamps (like many street lights) are more like gas discharge tubes in their operation.

- It is a good idea to emphasize that although the Bohr model of the atoms is important historically, it does not represent an accurate model for the atom. Bohr's model is no longer accepted because it is considered fundamentally incorrect (this is why the section is so short). A discussion of this model is useful in discussing the nature of science. The Bohr model is a relatively simple model with the sole intent of explaining the results of the hydrogen emission spectrum.

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[1] James E. Huheey, *Inorganic Chemistry; Principles of Structure and Reactivity*, 3rd ed., Harper and Row (1983).

[2] Linus Pauling, *The Nature of The Chemical Bond*, Cornell University Press (1960), pp. 511-519.

It succeeded in doing so and this is why it was considered powerful. However, this model does not fit polyelectronic atoms and electrons do not move in fixed orbits.

- In this section the idea of probability as a model for understanding nature is introduced. This will be new to all students and most will find it confusing. It is a fundamental change from determinism in looking at the world and it is understandable that students will have difficulty with it. For example, students will tend to think of orbitals as physical structures and will find it difficult to think of them as probability distributions. In Section 7.5 we describe the orbital as a *potential space* for an electron. This is a good way for the students to think about it initially.

- Many students confuse orbitals and orbits. An orbit describes a particular path an object follows as it travels around another. For example, the moon has an orbit about the earth. Electrons do not follow a particular path around the nucleus. An orbital describes the volume around the nucleus where an electron is likely to be found. The exact path of an electron in this area is not known.

- For polyelectronic atoms the third electron goes to the 2*s* orbital. In the hydrogen atom the 2*s* and 2*p* orbitals have the same energy. However, multiple electrons change this. With multiple electrons, in a given energy level the *s* orbitals are lower in energy (and thus filled first) than the *p* orbitals, which in turn are lower in energy than the *d* orbitals.

- Encourage the students to use the periodic table to assist them in figuring out electron configurations. Once they see that they can just move across the rows, filling electrons into the s and p orbitals, electron configurations become much easier to write. This should be used to emphasize the periodicity of the periodic table (even though the table was invented before the discovery of electrons). Thus, the students should realize that with an understanding of how the periodic table is put together, the students can figure out the expected electron configuration of any element. They need not memorize these. See Figure 7.28 in the text.

- Make sure students understand the trends for atomic size across rows and down groups in the periodic table. If they understand these trends, the trends for ionization energy are much easier to grasp. The students should be able to see that the trends for atomic size and ionization energy are consistent with one another.

- Many students believe that alkali metals “want” to lose an electron; that is, the potassium atom, for example, will release energy to become an ion. Or they believe that less energy is required to remove the second electron from calcium, for example, than the first because calcium “wants” to lose two electrons. Energy is always required to remove an electron, and successive electrons can only be removed with an increasing amount of energy.

CHAPTER EIGHT: BONDING: GENERAL CONCEPTS

Chapter Learning Goals:

Section One:	To explain why an ionic bond is formed. To explain why a covalent bond is formed. To introduce the polar covalent bond.
Section Two:	To discuss the nature of a bond in terms of electronegativity.
Section Three:	To define the relationship between bond polarity and molecular polarity.
Section Four:	To show how to predict the formulas of ionic compounds. To discuss the factors governing ion size.
Section Five:	To define lattice energy and show how it can be calculated.
Section Six:	To show the relationship between electronegativity and the ionic character of a bond.
Section Seven:	To discuss the covalent bonding model.
Section Eight:	To show how bond energies can be used to calculate heats of reaction.
Section Nine:	To introduce the localized electron model.
Section Ten:	To show how to write Lewis structures.
Section Eleven:	To show how to write Lewis structures for certain special cases.
Section Twelve:	To illustrate the concept of resonance. To show how to write resonance structures.
Section Thirteen:	To describe how molecular geometry can be predicted from the number of electron pairs.

Chapter 8 is the first of two chapters that covers chemical bonding. This arrangement is typical of most current general chemistry texts, but there are some important differences. One difference is the approach used in Chapter 8 and the second is the inclusion of VSEPR in this first chapter on bonding, rather than its placement in the chapter on orbitals.

The approach of Chapter 8 is to begin with a discussion of measurements that can be made that tell us something about the characteristics of bonds. Electronegativity is introduced so that measured dipole moments can be correlated to bond type. Ionic bonds are then discussed. The discussion of covalent bonding follows. The first section on covalent bonding begins with bond energies. Bond energies are, again, a quantity that can be measured. An emphasis is placed on the priority of measurement over theory, something that is lost in many books.

The use of VSEPR theory to predict molecular geometry is included in Chapter 8. We feel this is more logical for two reasons. First, molecular geometry is discussed in the same chapter as polarity. Second, this placement gives the instructor greater flexibility.

A key concept a student should get from the first-year course is to appreciate the intimate relationship between molecular structure and the bulk physical and chemical properties of matter. Including VSEPR in Chapter 8 places all of the topics necessary for dealing with this correlation in a single chapter. Instructors can then go directly to Chapter 10 and emphasize the structure-property relationships, leaving the more abstract topics on orbitals in bonding for a later time.

The section on formal charge should be considered optional. We feel that a discussion of formal charge is not essential in a first-year course and that there is a danger in overusing the concept. The section and end-of-chapter exercises are organized so that formal charge can be easily skipped. If you do choose to cover formal charge, we urge caution and suggest that you try to limit the discussion to using formal charge for deciding on reasonable resonance structures.

Teaching Tips

- Ionic and covalent bonding are the extremes of types of bonding. It is important that students understand that attractive forces between molecules lie on a continuum.

- Focus the student’s attention on where the electrons “spend more time”. Are they equally shared between two atoms or do they spend more time closer to one atom in the bond? This focus will help the students understand bond polarity.

- The relationship between the difference in electronegativity and bond type is rather arbitrary and only serves as a rough guide. Thus, the transition between polar covalent and ionic bonding, for example, is not a sharp one and the students should know this.

- Students will often confuse dipole moments in a molecule with the overall polarity of a molecule. Molecules with polar bonds may be non-polar molecules (CCl_4 is an example). The three-dimensional shape affects the overall polarity of a molecule (this is discussed in more detail in Section 8.13)

- The polarity of a molecule has a profound impact on its chemical and physical properties. For example, water is a liquid and methane (CH_4) is a gas at room conditions even though both have similar molar masses. Relating molecular shape to polarity to properties provides a “big picture” so the students know where we are going with this information.

- This is an excellent time to discuss the importance of polarity in molecules by using the properties of water. Discuss the ability of water to dissolve polar and ionic substances and the strong attraction that water molecules have for each other which results in a high boiling point for water.

- This is an excellent time to review the electron configurations for atoms and write electron configurations for ions. Emphasize that we can predict the ion that forms from a particular atom by achieving a noble gas electron configuration. The concept that there is a strong tendency for atoms to adopt a noble gas configuration is central to understanding the electronic structure of ions and the Lewis structures of molecules. This concept is emphasized in this chapter and should be stressed in class discussions. However, this does not explain *why* the atom achieves this configuration. Science is about observing patterns (laws) and developing explanations (theories or models). The fact that the atoms in a compound have a tendency to achieve a noble gas electron configuration is an observation, not an explanation. Knowing this helps us to make predictions, but it does not answer the “why?” question. Knowing this observation, students should be able to predict, for example, that a sodium atom will lose only one electron when becoming an ion, and that a calcium atom will lose two electrons. However, the reason is not because the atoms "want" to have a noble gas electron configuration. This is a subtle but important distinction.

  At the beginning of this discussion it should be emphasized that when chemists use the term *ionic compound*, they are usually referring to the solid state of that compound. In the solid state the ions are close together. That is, solid ionic compounds contain a large collection of positive and negative ions packed together in a way that minimizes the - .. - and +..+ repulsions and maximizes the +..- attractions. This situation stands in contrast to the gas phase of an ionic substance, where the ions are quite far apart on average. In the gas phase, a pair of ions may get close enough to interact, but large collections of ions do not exist. Thus, when we speak of the stability of an ionic compound, we are referring to the solid state, where the large attractive forces present among oppositely charge ions tend to stabilize (favor the formation of) the ions. For example, the O^{2-} ion is not stable as an isolated, gas-phase species but, of course, is very stable in many solid ionic compounds. That is, MgO(s), which contains Mg^{2+} and O^{2-} ions, is very stable, but the isolated, gas-phase ion pair $Mg^{2+} \bullet O^{2-}$ is not energetically favorable in comparison with the separate neutral gaseous atoms. Thus you should keep in mind that in this section, and in most other cases where we are describing the nature of ionic compounds, the discussion usually refers to the solid state, where many ions are simultaneously interacting.

- You can effectively show why noble gases do not generally form compounds by drawing the Lewis structures for several noble gases. However, it is a good idea to de-emphasize the drawing of Lewis structures for atoms. It is shown here to introduce simple Lewis structures (diatomic molecules in particular). But students are later confused when confronted with a molecule such as carbon monoxide (CO), for which we cannot draw the Lewis structure of the molecule directly from the Lewis structures of the atoms.

- Be sure to emphasize that in order to write correct structures for molecules it is necessary to sum all the valence electrons from all the atoms. Be sure that students do not attach valence electrons to individual atoms and then try to write a Lewis structure. This method can sometimes produce a correct structure but often does not for molecules containing multiple bonds and for complex molecules.

- Remind students that many elements do usually obey the octet rule. If an atom that can exceed the octet rule is bonded to atoms that obey the octet rule, the available electrons should be distributed to form octets on all the atoms. Then any remaining electrons can be placed on atoms that can exceed the octet rule.

- Use models in class to demonstrate each of the shapes in this section. It is difficult for students to visualize the shapes of molecules without models. Students can make models at home to help them with problems by using toothpicks and gumdrops or marshmallows. There is an activity in which the students make these shapes with balloons in the *Inquiry Based Learning Guide*.

- When teaching the VSEPR model, be sure to stress the difference between the arrangement of the electron pairs and the molecular structure. For example, point out the differences between $BeCl_2$ and water. This will help the students see the difference between electron pair structure and molecular shape. Many students do not understand why water should have a bent shape and $BeCl_2$ is linear.

- It is crucial that the students write the Lewis structure first. Many students believe they can predict the shape from the formula. However, it is easy to show counterexamples of this. For example, CO_2 is linear, while OF_2 is bent; BF_3 is trigonal planar, while NH_3 is trigonal pyramid.

CHAPTER NINE: COVALENT BONDING: ORBITALS

Chapter Learning Goals:

Section One:	To show how special atomic orbitals are formed in covalent bonding.
Section Two:	To show how molecular orbitals are formed in a molecule.

To define bond order and demonstrate how to calculate it.

Section Three: To discuss the bonding in certain molecules of the general formula X_2.
To relate paramagnetism to the filling of molecular orbitals.
To correlate bond order, bond energy, and bond length.

Section Four: To use the molecular orbital model to treat bonding between two different atoms.

Section Five: To show how the need for resonance is eliminated if the localized electron and molecular orbital models are combined.

Section Six: To show how photoelectron spectroscopy (PES) can be used to give information about the energies of electrons in molecules.

With the inclusion of geometry and VSEPR in Chapter 8, this chapter deals strictly with orbitals and bonding. Both valence bond and molecular orbital models are discussed. Instructors who wish, either because of time constraints or philosophy, to do so can delay this chapter to later in the second semester or use parts as appropriate.

The end-of-chapter exercises try to emphasize some of the more practical applications of the bonding models. For example, consideration of the nature of the π-bonds leads one to the conclusion that C_2H_4 is planar and C_3H_4 is not. The exercises on M.O. theory also emphasize the relationship between the model (predicted bond orders) and observations (bond energies, bond lengths, and paramagnetism).

An additional topic instructors may wish to introduce at this point is the difference between second period elements and their heavier congeners. Much of the difference between Si and C (Chapter 20), for example, can be attributed to differences in bonding.

Teaching Tips

- The idea of hybridization and molecular orbitals highlights the concept that molecules are not simply “atoms stuck together”. The atomic orbitals change to accommodate the formation of molecules.

- A discussion of hybridization is another area in which you can discuss the nature of science. In this case, we create a model based on empirical evidence (observations). The fact that the bond angles in methane, for example, are not 90° leads us to think about how atomic orbitals must change in the formation of compounds.

- Figures 9.5, 9.9, and 9.16 show the students to use “lowest energy arguments” when determining the stability of a molecule. This complements the discussion that began in Chapter 6.

- Point out Figure 9.24 to students. This is a good summary to show the students that hybridization is dependent on geometry, which comes from the Lewis structures. So, if you ask a student to determine the hybridization of a certain atom in a molecule, they are not expected to just “know” it – they must write out the Lewis structure first.

- Some students will find it very helpful to make physical models of figures such as Figure 9.13 and Figure 9.19.
- The molecular orbital model relies on the fact that orbitals are wave functions and that waves can constructively interfere or destructively interfere. Thus, we get bonding or antibonding orbitals, respectively.
- Use Figure 9.38 to show how bond order (from MO theory) is related to the number of bonds in a Lewis structure. This shows that the models are consistent.
- Figure 9.26 (and others like it) shows not only the molecular orbitals, but the atomic orbitals to show relative energies. Figure 9.38 shows only the molecular orbitals.

CHAPTER TEN: LIQUIDS AND SOLIDS

Chapter Learning Goals:

Section One:	To define dipole-dipole force, hydrogen bonding forces, and London dispersion forces.
	To describe the effects these forces have on the properties of liquids and solids.
Section Two:	To describe some properties of liquids: surface tension, capillary action, and viscosity.
Section Three:	To contrast crystalline and amorphous solids.
	To introduce X-ray diffraction as a means for structure determination.
Section Four:	To discuss the concept of closest packing of metal atoms.
	To describe two models for bonding in metals.
	To define and classify alloys.
Section Five:	To show how the bonding in elemental carbon and silicon accounts for the widely different properties of their compounds.
	To explain how a semiconductor works.
Section Six:	To describe the bonding in molecular solids.
Section Seven:	To model the structures of ionic solids using the packing of spheres.
Section Eight:	To define the vapor pressure of a liquid.
	To discuss the features of heating curves.
Section Nine:	To discuss the features of phase diagrams.

Chapter 10 is one of the key chapters in the text. The correlation of structure and properties is one of the thought processes central to all of chemistry. Solid-state chemistry is becoming increasingly important.

One of the key thought processes used by chemists and one that is perhaps unique to chemists, the correlation between structure and properties, is the idea that permeates this chapter. From the nature of the bonding and the geometry of molecules, we can predict the nature and relative strengths of interparticle forces. With that in hand we can begin to either predict or correlate physical properties to structure. The same type of reasoning can be extended to rationalizing and categorizing chemical properties.

Topics such as types of crystals, unit cells, closest packing, vapor pressures, and phase diagrams can be given variable emphasis depending on the students' and instructor's interests. With a class of mostly engineers, for example, extra emphasis might be given to the solid state and particularly semiconductors. Or a discussion of metallurgy (Chapter 21) can be covered here.

Steelmaking and polymer additives (Chapter 21 and 22) are excellent examples of how the properties of materials are chemically fine-tuned. If the extensive section on polymer additives is discussed at this point, at least a brief discussion of organic compounds and nomenclature (Chapter 22) would be in order. Alternatively, the section on the differences between carbon and silicon can be expanded by dealing more with their compounds here (Chapters 20 and 22).

For example, at the University of Illinois at Urbana-Champaign, one of the second-semester general chemistry courses (Chemistry 104) is aimed primarily at engineering students. We begin the second semester by covering Section 10.1 on intermolecular forces and then spend about three weeks on organic chemistry and synthetic polymers, emphasizing structure-property relationships. Later we come back to the rest of Chapter 10 when we talk about condensed phases, placing an emphasis on the solid state and electrical properties of solids. This chapter is really a pivotal chapter in the book. The overall flavor of a course is greatly affected by how instructors choose applications to illustrate the topics in this chapter.

Teaching Tips

- Students may need to be reminded that energy is required to change phase even though the temperature is not changing during that time. The energy applied to the system is used to overcome the intermolecular forces.

- Remind the students that a phase change occurs when no temperature change occurs and no change of phase occurs when the temperature is changing.

- Be sure to help the students make a clear distinction between intramolecular forces (covalent and ionic bonds) and intermolecular forces.

- This is an excellent opportunity to review the main concepts from bonding and prepare the students to look at the forces that exist between molecules. It is especially helpful to review the concept of a dipole and what it means before moving into intermolecular forces. This is the time to tie in electronegativity and arrangement of atoms to polarity of molecules.

- Some students think that dipole-dipole interactions and hydrogen bonds are completely unrelated. Be sure to stress that all of these interactions result from polarity.

- Students are often confused by the term hydrogen bond. They have the misconception that the hydrogen bond is the O-H, N-H, or F-H bond and not the attractive force between the dipoles. We have used dashed lines to represent hydrogen bonds to help the students see this more clearly. Stress that a hydrogen bond is an attractive force between molecules.

- Students have trouble deciding when molecules should have hydrogen bonds between them. Use examples such as NH_3, CH_3OH and HF to show that in each case the hydrogen is bonded directly to the N, O, or F. Molecules such as $(CH_3)_3N$ do not form hydrogen bonds even though they contain both N and H because there is no hydrogen bonded to the N to produce the dipole needed for a hydrogen bond.

- It is hard for student to visualize what happens within a molecule to produce London dispersion forces. This is difficult to represent with a picture. It is easiest to begin talking about these forces by using small nonpolar molecules.

- London dispersion forces exist not only among noble gas atoms and nonpolar molecules, but also between all molecules. However, they are generally only important among noble gas atoms and nonpolar molecules because they are relatively weak.

- The idea that evaporation is an endothermic process is hard for many students to grasp. Be sure to take time to explain that the temperature of the liquid drops because the higher kinetic energy molecules are going to the gas phase. Energy flows in if the container is not insulated and the temperature of the liquid will then remain constant.

- The familiar scents of finger nail polish remover and rubbing alcohol tell us that these liquids evaporate relatively easily. They have relatively high vapor pressures.

- Vapor pressure is really the beginning of introducing the concept of equilibrium to students. The idea that molecules are changing between the gas and liquid phase at constant equal rates is new for students. Focus the attention of the students on Figure 10.38 to help them understand what is taking place at the molecular level.

- Remind students that the formulas for ionic solids are actually empirical formulas, not molecular formulas.

CHAPTER ELEVEN: PROPERTIES OF SOLUTIONS

Chapter Learning Goals:

Section One:	To define various ways of describing solution composition.
Section Two:	To define the heat of solution and discuss its various energy components.
Section Three:	To show how molecular structure, pressure, and temperature affect solubility.
Section Four:	To show how a solution's vapor pressure is affected by the concentration of solute and the interactions of solute and solvent.
Section Five:	To explain the effect of a solute on the boiling and freezing points of a solvent.
Section Six:	To explain osmosis and describe its application.
Section Seven:	To show how the colligative properties of electrolyte solutions can be used to characterize the solute.
Section Eight:	To define a colloid and explain how it is stabilized.

The treatment of solutions is fairly standard. An emphasis can be placed on structure-property correlation here, too. Enthalpy changes and thermochemical cycles can be discussed further by dealing with the energetics of solution formation. The remaining topics can be given whatever emphasis the instructor chooses.

Instructors who do not choose to do all of Chapter 4 early can begin introducing some of these topics (composition, nature of water, precipitation reactions) at this point in the course. It would also work well to go in the following order: Sections 4.1 to 4.3, Chapter 11, Sections 4.4 and 4.12. If Chapter 4 was covered in entirety early in the course, students can use Problems 1-8 in the end-of-chapter exercises as a review.

Teaching Tips

- A common example of a solution with a concentration expressed in mass percent is vinegar. You may want to bring a bottle to class showing a label giving the mass percent (usually 5%). You can have the students convert this to molarity, given the density.

- This is an excellent opportunity to repeat the conductivity demonstration used in earlier chapters. This will remind students that we can tell if a solution contains ions by it conductivity.

- Students can remember “like dissolves like” but often have trouble figuring out which substances are alike and which are not. Use the structures of similar molecules such as CH_3OH and water to show their similarities.

- Explaining why two non-polar substances mix is difficult without bringing in the concept of entropy. The text addresses this with a discussion of probability of mixing (see page 503).

- Although the text uses the term *hydrophobic*, the term is actually a misnomer. While a substance such as oil does not mix with water, a nonpolar molecule does not repel a polar molecule. In fact, they have attraction for one another (London dispersion forces). However, the attraction for a polar molecule with another polar molecule is greater, and it looks as though oil and water, for example, repel one another when we mix them together.

- Henry’s law and the discussion of vapor pressure introduce students to the idea of equilibrium (as did Section 10.8). This can be recalled when discussing chemical equilibrium in Chapter 13.

- The concentration unit molality is used for colligative properties because it is not dependent on a change in temperature (like molarity is).

- Note that the osmotic pressure formula, $\pi = CRT$ is analogous to the ideal gas law formula $P = (n/V)RT$. It turns out this stems from the model that vapor pressures are important when considering why a solution exhibits osmotic pressure.

CHAPTER TWELVE: CHEMICAL KINETICS

Chapter Learning Goals:

Section One:	To define the reaction rate and to show how rates can be measured from experimental data.
Section Two:	To describe the two types of rate laws.
Section Three:	To learn methods for determining the rate law for a reaction.
Section Four:	To develop rate laws relating concentration to reaction time and to show how they can be used to determine reaction order.
Section Five:	To explore the relationship between the reaction pathway and the rate law.
Section Six:	To discuss the temperature dependence of reaction rates. To describe the collision model. To define and show how to calculate activation energy.
Section Seven:	To explain how a catalyst speeds up a reaction. To discuss heterogeneous and homogeneous catalysis.

Chemical kinetics seems to be one of the most difficult topics for students. Some of the topics, such as collision theory, are abstract -- one of the reasons we have included an increased emphasis on pictorial representations of reactions in this and other chapters. Other topics, such as rate laws and graphing, require a good math background, a great deficiency in many beginning students.

Although this chapter is written to make this material understandable to students at this point in the course, some instructors may wish to postpone this chapter until after all of the equilibrium chapters are covered. Chapters 12 and 13 can also be reversed in order, as a full treatment of kinetics is not required before covering Chapter 13. Equilibrium is discussed from a phenomenological standpoint rather than on a kinetic basis. Another alternative order is to do kinetics with the exceptions of mechanisms first, then take up equilibrium, and finish with mechanisms. This approach allows one to cover in greater detail mechanisms that involve an equilibrium step. Instructors who covered nuclear reactions (Chapter 19) early with Chapter 2 can use the kinetics of nuclear decay as an example of a first-order reaction. Some of the end-of-chapter exercises cover nuclear decay. This allows for the mention of nuclear decay by instructors who do not cover Chapter 19 in detail.

Because we have placed kinetics before equilibrium, we have eliminated all end-of-chapter exercises that concern mechanisms with a fast equilibrium before the rate-determining step. Many instructors may wish to switch the order of these chapters.

Teaching Tips

- The study of chemical kinetics allows us to think about *how* chemical reactions occur. Stoichiometry tells us what happens (and amounts) and thermodynamics helps us to predict what will happen.

- Deriving the integrated rate law equations requires calculus.

- Watch for the misconception that a reaction is complete after two half-lives (in fact, the reaction is 75% complete after the first two half-lives). Use Figure 12.X when discussing this.

- Table 12.6 nicely summarizes the discussions of differential rate laws, integreated rate laws, and half-lives.

- Use Figure 12.10 to illustrate the energy changes that occur when 2 BrNO molecules from Figure 12.12 interact.

- Most collisions do not result in a chemical reaction. This is due to energy requirements and positioning of the molecules. If all collisions resulted in a chemical reaction, life would be extremely unstable. Virtually everything would happen at a much higher speed (including our aging). Students should appreciate the vast number and types of chemical reactions that are occurring all around them (especially in an atmosphere rich in oxygen).

- Mechanisms cannot be proven, but the observed rate laws can be used to support them.

- Remind students that the kinetic energy of a collection of molecules is directly proportional to the temperature (K).

- To get to Equation 12.11 in the text we must assume that A is not dependent on temperature. However, the value of A must be dependent on temperature since part of it is due to the number of molecular collisions. However, for relatively small temperature differences, this equation works (this is true because the number of collisions is relative to the square root of the temperature, and since must use a Kelvin scale, small differences in temperatures do not appreciably change the value of A.

- This is an excellent time to discuss the use of catalysts. Catalytic converters in automobiles use metal such as palladium and platinum to convert uncombusted hydrocarbons to carbon dioxide and water. The metals are used repeatedly in these processes. In the human body enzymes act as catalysts speeding up reactions that would take hours in the laboratory to fractions of a second in the body.

CHAPTER THIRTEEN: CHEMICAL EQUILIBRIUM

Chapter Learning Goals:

Section One:	To discuss how equilibrium is established.
Section Two:	To introduce the law of mass action and to show how to calculate values for the equilibrium constant.
Section Three:	To show how K and K_p are related.
Section Four:	To show how condensed phases are treated in constructing the equilibrium expression.
Section Five:	To show how the equilibrium constant is used to predict the direction in which a system will move to reach equilibrium. To demonstrate the calculation of equilibrium concentrations given initial concentrations.
Section Six:	To generalize the procedure for doing equilibrium calculations.
Section Seven:	To show how to predict the changes that occur when a system at equilibrium is disturbed.

Students generally find the topic of chemical equilibrium to be quite difficult. One difficulty is that we treat reactions as macroscopically static (the concentrations do not change over time once equilibrium is reached) yet microscopically dynamic. Students do not always differentiate between the two, and usually visualize the reactions as static (the macroscopic is given precedence over the microscopic). This is why it is important to emphasize representations of the reactions, which is done in the text and asked of the students in some of the Active Learning Questions.

The chapter on chemical equilibrium is written to be a stand-alone chapter. It does not require that kinetics be covered first. There are exercises at the end of the chapter that cover the interplay between equilibrium and kinetics, and these can be assigned after both topics have been covered. The chapter introduces, primarily through gas phase examples, the phenomenon of equilibrium and the techniques that are useful in solving equilibrium problems. Appendix 1.4, Solving Quadratic Equations, can also be covered here. It describes the method of successive approximations that can be very useful in solving the equilibrium problems at the end of the chapter. Other examples of gas-phase reactions from other chapters can be brought in here to cover more descriptive topics. The synthesis of ammonia and the chemistry of nitrogen oxides (Chapter 20) or the petrochemical chemistry (Chapter 22) can be used as illustrations of gas-phase reactions.

Teaching Tips

- Point out that equilibrium can only be reached in closed systems. This includes gases in closed containers and ions in solution (since the ions cannot escape). In industry processes do not usually reach equilibrium. Instead the process is designed to maximize the amount of product formed (and avoid the limitation of equilibrium) by removing the product as it is formed.

- Emphasize that equilibrium in not a condition in which the reactants and products are present in equal amounts. The balance is between the rates of the forward and reverse reactions.

- Students often think that equilibrium is a static condition. Be sure to emphasize that, although at the macroscopic level things appear to be constant at equilibrium, at the molecular level chemical reactions continue to take place. The rates of the forward and reverse reactions are equal.

- Be sure to use balanced chemical equations including states for each chemical species in this section. In the next section we discuss heterogeneous equilibria and the students will need to distinguish among the states of the chemicals.

- Once students begin to understand how to setup the equilibrium expression for a reaction we can begin to talk about calculations involving the expression. It is important to distinguish between initial conditions and equilibrium conditions when students begin calculations.

- Have the students calculate the change in concentrations for each of the species in Table 13.1. Make sure the students understand that the ratio of concentration changes is the same ratio as the coefficients in the balanced equation.

- Use the results in Table 13.1 to emphasize the difference between "equilibrium position" and "equilibrium constant."

- When calculating K from concentrations of reactants and products, it is important to emphasize both initial and equilibrium concentrations so that the students can distinguish between the two sets of data. This can be a major source of confusion for beginning students when working these problems.

- Students have trouble understanding how a change in volume affects the system. The idea that decreasing the volume increases the pressure and therefore changes the concentration is difficult for many students to understand.

- When working with gases it is important for the students to remember that they must use a balanced equation to make their predictions.

- For students who would like a mathematical verification of the effect of a change in volume, discuss how the partial pressures of the gases are changed. For example, suppose that at some temperature the equilibrium concentrations of N_2, H_2, and NH_3 are all 1.0 *M* (thus, at this temperature $K = 1.0$). If we suddenly double the volume, the concentrations of each species initially decrease to 0.50 *M*. If we put these numbers into the equilibrium expression we get a value of 4.0.

This number is higher than *K* so the system must shift to the left to satisfy the equilibrium condition (K = 1.0). This is the direction predicted by LeChatelier's principle.

CHAPTER FOURTEEN: ACIDS AND BASES

Chapter Learning Goals:

Section One:	To discuss two models of acids and bases and to relate equilibrium concepts to acid dissociation.
Section Two:	To relate acid strength to the position of the dissociation equilibrium. To discuss the autoionization of water.
Section Three:	To define pH, pOH, and pK and to introduce general methods for solving acid-base problems.
Section Four:	To demonstrate the systematic treatment of solutions of strong acids.
Section Five:	To demonstrate the systematic treatment of solutions of weak acids. To show how to calculate percent dissociation.
Section Six:	To introduce equilibria involving strong and weak bases. To show how to calculate pH for basic solutions.
Section Seven:	To describe the dissociation equilibria of acids with more than one acidic proton.
Section Eight:	To explain why certain salts give acidic or basic solutions and to show how to calculate the pH of these solutions.
Section Nine:	To show how bond strength and polarity affect acid-base properties.
Section Ten:	To show how to predict whether an oxide will produce an acidic or basic solution.
Section Eleven:	To define acids and bases in terms of electron pairs.
Section Twelve:	To summarize the major species approach to solving acid-base problems.

This chapter emphasizes a problem-solving approach to acid-base chemistry first used in *General Chemistry Problem Solving II*, by Steven S. Zumdahl (D. C. Heath, 1979).[3] The key features of the approach are to recognize which acids and bases are present, what reactions are possible, and how to break complicated problems down into more manageable ones.

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[3] These problem-solving techniques are also covered in *Solving Equilibrium Problems with Applications to Qualitative Analysis*, D. C. Heath, 1988. The chapters on equilibrium are the same in both of these books.

In many instances the general chemistry course is the only exposure to aqueous equilibria that students will get. As a result, it is a topic that is covered fairly rigorously. The combination of chemistry with mathematical rigor gives many students a great deal of difficulty. Developing the students' ability to think things through in a systematic fashion will get them through acid-base equilibria with lessened anxieties and frustrations.

The text outlines a systematic approach for solving equilibrium problems. It consists of the 11 steps outlined in the text. For a weak acid the first three steps are:

1. List the major species in the solution.
2. Choose the species that can produce H^+, and write balanced reactions for the reaction producing H^+.
3. Using the values of the equilibrium constants you have written, decide which equilibrium will dominate in producing H^+.

In our experience these three steps are the key considerations and should be strongly emphasized. Listing major species gives the student a method for attacking these problems that emphasizes thinking about chemistry rather than memorizing examples. This approach works for all problems involving solution equilibria. Thus the student can look within the problem for the solution rather than outside (memorized pigeonhole examples that usually don't apply to the actual problem at hand anyway). The early introduction to acid-base reactions in Chapter 4 starts the student off with the importance of thinking about reactions. The other important skill is to recognize structural features that give rise to acidic or basic properties. This idea again amplifies the importance of the relationship between structure and properties first covered in Chapter 10. Once the acid (or base) has been identified, a base (or acid) must be found in the species in solution (most often water) that can accept (or donate) a proton. Then it is easy to write a reaction, realizing that acid-base reactions involve a transfer of protons.

The next step is to get values of equilibrium constants for the reaction. This requirement causes difficulty for many students when they must calculate an equilibrium constant from a tabulated one. For example, for the reaction

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

only K_b for ammonia is given; the student must calculate K_a from

$$K_aK_b = K_w$$

In many textbooks much of the confusion results from calling the above reaction a hydrolysis reaction and giving the equilibrium constant the symbol, K_h. Students think that this is another kind of reaction that must be treated as something different; they get confused and frustrated. We have avoided the use of the term "hydrolysis." There are acids and there are bases. Ammonium ion is an acid because H_3O^+ is being produced.

If there are only acids and bases, it is easier for the student to see that it doesn't make sense to use an equilibrium constant that pertains to a base for the dissociation of NH_4^+. The student is directed to focus attention on a relationship between NH_4^+ and NH_3, the conjugate acid-base pair. The thought naturally arises to look for a relationship between K_a and K_b for a conjugate acid-base pair. The text emphasizes thinking in a logical fashion, looking for structural features, and writing the reactions when confronted with acid-base equilibria. This strategy will serve the student much better than attempting to pigeonhole every type of problem.

In the solving of the resulting quadratic equation, the text develops several methods of simplifying the algebra. Assumptions are used to simplify the math when possible, and the method of successive approximations is presented as an approach to use when the usual assumptions are invalid. The method of successive approximation is explained more fully in an appendix and used extensively in the *Solutions Guide*. This tactic further emphasizes the goals of the chapter. We want to deal with the chemistry of acids and bases; we want to make the algebra simple.

The supplements *General Chemistry Problem Solving II* and *Solving Equilibrium Problems with Applications to Qualitative Analysis*, both by Steven S. Zumdahl, can also be used during the coverage of aqueous equilibria. They contain the same systematic approach used to solve equilibrium problems in the main text and provide a great number of simpler, straightforward acid-base and buffer problems. In addition, each chapter contains several multiple-choice questions that the student can use in a self-learning drill type fashion.

Teaching Tips

- Spend time developing the model for acid-base pairs in this section. Students need to be able to look at a chemical equation and identify the acid-base pairs.

- The Bronsted-Lowry definition of acids and bases is preferable because it emphasizes to students that for one reactant to act like an acid the other reactant must act like a base.

- It is important in this section to show students that these substances work together. A proton donor must have a substance to accept the proton. The idea that water can act as both an acid and as a base can be developed here.

- This is the first time students have seen the hydronium ion. Be sure that they understand how it forms from the donation of a proton to water.

- Have students draw molecular level pictures to illustrate the difference between strong acids and weak acids. This idea is important to the understanding of how acids behave.

- We use the terms “ionized” and “dissociated” interchangeably when discussing acids.

- For now, emphasize the competition for a proton between a conjugate base and water in determining which reaction (forward or reverse) dominates.

- After discussing forward and reverse reactions, the students should understand that $C_2H_3O_2^-$ is a better base than H_2O, and H_2O is a better base than Cl^-.

- In discussing the phrase “A weak acid contains a relatively strong conjugate base”, emphasize the term “relatively”. The conjugate base of a weak acid is not a strong base in the way that NaOH is a strong base. This is because for the reaction of the conjugate base in water, the reverse reaction dominates. However, if we compare two unequally weak acids, the conjugate base of the weaker of the two acids is a stronger base, although both bases are considered weak (compared to OH^-).

- Use an equation with two water molecules on the reactant side to show the students that water acts as an acid and as a base. In this way the students can see that a proton is donated from one water molecule and it is accepted by the other water molecule. This allows them to correctly identify acid-base pairs.

- While the students have experience with the concept of an equilibrium constant, they may still not apply it well here. When you discuss the idea that K_w is a constant, be prepared for students who reason “Pure water contains 1.0×10^{-7} *M* H^+ and 1.0×10^{-7} *M* OH^-. If we add 1.0 mol of HCl to a liter of water, the concentration of H^+ is about 1.0 *M*. So isn't $K_w = [H^+][OH^-] = [1.0\ M][1.0 \times 10^{-7}\ M] = 1.0 \times 10^{-7}$?” These students do not understand that added H^+ will react with OH^-, thus lowering the concentration of OH^-. If 1.0 mol of H^+ is added to a liter of water, the OH^- concentration would be 1.0×10^{-14} *M*. This will not be obvious to the students at this point.

- It is helpful for students to go through all equilibrium equations of a polyprotic acid at least once to prove to themselves that only the first dissociation generally matters for the pH.

- Emphasize the definitions of acidic, basic, and neutral solutions. Notice however that the emphasis should be on the relationship of the concentrations of the ions, not on the pH of the solution. This reinforces the concept of *major species* that is highlighted in the text.

- Salt solutions are not new since they consist of conjugate acids or bases in solution. Students may try to treat them differently from weak acids or weak bases, but emphasize the similarities.

CHAPTER FIFTEEN: APPLICATIONS OF AQUEOUS EQUILIBRIA

Chapter Learning Goals:

Section One:	To study the effect of a common ion on acid dissociation equilibria.
Section Two:	To explain the characteristics of buffered solutions. To show how to calculate a buffer pH given the concentrations of the buffering materials.
Section Three:	To describe the meaning of buffer capacity.
Section Four:	To demonstrate how to calculate the pH at any point in an acid-base titration.
Section Five:	To explain how acid-base indicators work.

This chapter is essentially a continuation of Chapter 14. It begins with more complicated acid-base topics: buffers, titrations, and indicators. The problem-solving approach to acid-base problems is continued in this chapter. It helps to emphasize that these problems respond to the same techniques introduced in Chapter 14: writing major species and breaking down problems into smaller parts. An important example of this type of logical thought comes in solving problems that involve complex ion formation. From the values of equilibrium constants, the student knows that the reaction goes almost to completion. Breaking the overall problem down into a stoichiometry problem (complete reaction) followed by an equilibrium problem makes the algebra easier. The chapter emphasizes thinking through the chemistry before doing the algebra. Explicitly calling the first step in the process a stoichiometry problem lets the students see they are not doing anything new. It helps them see why we sometimes approach equilibrium in two steps rather than one. If this explicit label is not provided, it has been our experience that this approach can confuse students. They regard the reaction first going to completion as a mathematical trick rather than chemistry that they have already seen.

Teaching Tips

- A discussion of buffered solutions requires an understanding of LeChatelier’s Principle.

- The concept of a buffer is quite complex. It is important to go slowly through the explanation of how a buffer works. Students need to see each of the possibilities and what happens.

- Discuss the chemical properties that are necessary to produce a buffer. Ask questions such as: Why is important to use a weak acid and its salt? Why can’t we use a strong acid and its salt?

- The Henderson-Hasselbalch equation assumes that the vale of x is negligible in an ICE table. It is not necessarily wrong to assume this (in fact, for most buffers it is fine), but students should always be aware of the assumptions they make.

- Titrations are not “new” but they are good summaries of all of the chemistry that the students have learned in Chapters 14 and 15. Point out Figure 15.3 (the titration of a weak acid with a strong base) to show the students where the solution is buffered, for example.

- Figure 15.4 is a good one to show students that the equivalent point is not dependent on the type of acid or base, but their amounts.

CHAPTER SIXTEEN: SOLUBILITY AND COMPLEX ION EQUILIBRIA

Chapter Learning Goals:

Section One:	To show how to calculate the solubility product of a salt given its solubility, and vice versa. To demonstrate the prediction of relative solubilities from K_{sp} values. To explain the effect of pH and a common ion on the solubility of a salt.
Section Two:	To show how to predict whether precipitation will occur when solutions are mixed. To describe the use of selective precipitation to separate a mixture of ions in solution.
Section Three:	To apply the principles of equilibrium to the formation of complex ions. To show how complex ion formation can increase the solubility of a salt.

This chapter provides the first exposure to complex ion formation. Instructors can integrate a discussion of transition metal chemistry (Chapter 21) with the discussion of equilibria.

Some texts treat solubility equilibria before acids and bases. This is a reasonable alternative. It is probably done because for simple salts like AgCl the algebra is easier than that encountered in acid and base problems. This text has not taken such an approach because the acid-base chemistry thoroughly introduces the use of approximations in solving equations, and the solubility of some salts also depends on both acid-base and complex ion equilibria. The foundation is laid in the acid-base sections; all aspects of solubility and complex ion formation can be built upon that foundation. In addition, using only solubility product relationships tends to oversimplify a lot of what goes on in solution. For example, even in the case of silver chloride, a significant amount of the material in aqueous solution exists as neutral AgCl. People not wishing to tread down this path can rigorously treat acid-base equilibria and then deal with solubility and complex ions in a qualitative fashion.

Teaching Tips

- Students often think of solubility product constants as somehow different from equilibrium constants in Chapter 13 (presumably because of the solids in the chemical equations). Point out the the concept of equilibrium is the same in each case.

- Relative solubilities of compounds and their K_{sp} values are only related if there are the same number of ions in the compounds. If there are a different number of ions, we cannot say something like "Because the K_{sp} value of compound A is lower than that of compound B, compound A must be less soluble than compound B." This claim is true if there is the same number of ions in each compound.

- Have students prove to themselves that if we add chemical equations to get an overall equation we multiply the values of their equilibrium constants to get an overall equilibrium constant.

CHAPTER SEVENTEEN: SPONTANEITY, ENTROPY, AND FREE ENERGY

Chapter Learning Goals:

Section One:	To define a spontaneous process. To define entropy in terms of positional probability.
Section Two:	To state the second law of thermodynamics in terms of entropy.
Section Three:	To discuss the important characteristics of entropy changes in the surroundings. To apply the relationship between ΔS_{surr}, ΔH, and T (K).
Section Four:	To define free energy and relate it to spontaneity.
Section Five:	To apply positional probability to chemical reactions. To relate molecular complexity to entropy.
Section Six:	To show how to calculate the standard free energy change in a chemical reaction. To define standard free energy of formation and show how to use it to predict spontaneity.
Section Seven:	To relate free energy to pressure.
Section Eight:	To define equilibrium in terms of minimum free energy. To show how the value of K is related to ΔG°.
Section Nine:	To relate work done to the change in free energy.

The discussion of thermodynamics has been split into two chapters. Chapter 6, Thermochemistry, deals essentially with the first law and conservation of energy. It is placed early so students can get an early exposure to energy relationships, and it gives the background necessary for dealing with the energetics of bonding and phase transitions. Some of this chapter could be discussed at that earlier point. Alternatively, Chapter 6 could be delayed until just before Chapter 17.

Our preference is to keep the chapters split. The students need an early introduction to energy and thermochemistry, but they might be overwhelmed with too much thermodynamics all at once. The later placing of the second law accomplishes this goal. The chapter also places the emphasis of the second law on entropy. This focus would be lost by doing the second law early. The emphasis would tend to shift to free energy.

A focus on entropy, in our opinion, makes the second law clearer. The idea of the formation of stronger bonds (ΔH) serving as a driving force for chemical change can be developed from the topics in Chapter 6. As the intervening chapters are covered, the students become aware that another driving force must be present. Many endothermic processes occur spontaneously. With this appreciation built up over a significant length of course time, it is easier to introduce entropy. Good analogies can be used to qualitatively develop the concept of randomness as a driving force. For example, work has to be done to arrange every student in class in alphabetic order and it work has to be performed to cheat at cards to get a better hand. Chapter 17 is fairly abstract, but it is quite pictorial in approach. We have found that students can develop a better intuitive feel for randomness (entropy) as a driving force for change than they do for free energy. Once the focus of the chapter is placed on entropy, it is easier to accept free energy as a function that accounts for both driving forces for chemical change.

Teaching Tips

- The focus of the section is to introduce the concept that energy naturally changes from useful forms to less useful forms. Each energy transformation results in the same amount of energy but changed to a less useful form. The focus on the quality of energy rather than the quantity of energy is the key idea to present in this chapter.

- The first law of thermodynamics tells us that energy is conserved. The second law gives us some indication of the direction of this energy (the second law is often described as “time’s arrow”). For example, the first law cannot explain why we could not “uncook” an egg by putting a cooked egg in the refrigerator; or why a ball which has just rolled down a hill (and has transferred energy to the hill) could not roll up the hill by spontaneously gathering the newly transferred energy from the hill. The second law informs us that these phenomena are impossible.

- Physical disorder is often given to get the students to think about entropy. Actually, it is best to think of entropy in terms of energy states. However, this is rather abstract. One way to discuss this is to remind students that molecules have average energies, and temperature is a measure of average kinetic energy. Use Figure 5.21 on page 212 to show students how there is more variance in the "energy spread" of a gas at higher temperatures.

- Students often misinterpret the second law of thermodynamics by thinking that the entropy of the system must increase. It is the entropy of the universe that must increase.

- It is often the case that entropy and enthalpy run counter to each other in the direction of spontaneity. For example, ice requires energy to melt (it is endothermic, and thus the enthalpy term could be said to be unfavorable for spontaneity). However, water is more disordered than ice, so entropy is favorable. Which factor "wins"? It depends on temperature.

CHAPTER EIGHTEEN: ELECTROCHEMISTRY

Chapter Learning Goals:

Section One:	To describe the half reaction method for balancing oxidation reduction reactions.
Section Two:	To review oxidation and reduction. To define the components of an electrochemical cell. To distinguish between a galvanic cell and an electrolytic cell. To define cell potentials.
Section Three:	To describe how standard reduction potentials are assigned in terms of the standard hydrogen electrode. To demonstrate the combination of half-reactions to form the cell reaction. To characterize a galvanic cell.
Section Four:	To relate the maximum cell potential to the free energy difference between cell reactants and products.
Section Five:	To discuss the driving force in concentration cells. To quantify how to calculate the relationship between cell potential and cell concentration. To show how to calculate equilibrium constants from cell potentials.
Section Six:	To discuss the composition and operation of commonly used batteries.
Section Seven:	To explain the electrochemical nature of corrosion and describe some means for preventing it.

Section Eight:	To describe the stoichiometry of electrolysis reactions. To show how to predict the order of electrolysis of the components of a mixture.
Section Nine:	To discuss the manufacture of aluminum, the chlor-alkali process, and other industrial applications of electrolysis.

The chapter begins with a review of oxidation-reduction reactions (first discussed in Chapter 4). The emphasis is on using the half reaction method to balance oxidation-reduction reactions.

Free energy is used to develop some of the principles in the electrochemistry chapter, and Chapter 18 follows naturally after Chapter 17. As a pair, these chapters could be placed at different places in the course, as previously mentioned for Chapter 17. Instructors who have not previously covered oxidation-reduction in Chapter 4 need to cover the pertinent topics from Chapter 4 before Chapter 18. At the beginning of the end-of-chapter exercises in Chapter 18, there are problems to review the oxidation-reduction topics covered in Chapter 4.

One problem that students have in electrochemistry is dealing with sign conventions for half-cell potentials, and understanding how those half-cell potentials are combined to get an overall cell potential. This process is simplified in this text by emphasizing that, to get an oxidation-reduction reaction from two reduction half-reactions, one of the half-reactions must be reversed (and the sign of the E° value reversed). The decision of which one to reverse is based on getting a net positive potential for the cell. For example, for the reaction

$$\mathrm{Cu(s)} + 2\,\mathrm{Ag^+(aq)} \rightarrow \mathrm{Cu^{2+}(aq)} + 2\,\mathrm{Ag(s)}$$

we calculate E° in the following manner. Two half-reactions are

$$\begin{array}{llllll} \mathrm{Ag^+} & + & \mathrm{e^-} & \rightarrow & \mathrm{Ag} & E^\circ = 0.80\ \mathrm{V} \\ \mathrm{Cu^{2+}} & + & 2\,\mathrm{e^-} & \rightarrow & \mathrm{Cu} & E^\circ = 0.34\ \mathrm{V} \end{array}$$

To get a positive E°, the Ag/$\mathrm{Ag^+}$ half-cell involves the reduction (cathode) and the Cu/$\mathrm{Cu^{2+}}$ involves oxidation (anode). Thus

$$\begin{array}{llll} \text{reduction} & 2(\mathrm{Ag^+} + \mathrm{e^-} & \rightarrow \mathrm{Ag}) & E^\circ = 0.80\ \mathrm{V} \\ \text{oxidation} & \quad \mathrm{Cu} & \rightarrow \mathrm{Cu^{2+}} + 2\,\mathrm{e^-} & E^\circ = -0.34\ \mathrm{V} \\ \hline & \mathrm{Ag^+} + \mathrm{Cu} & \rightarrow \mathrm{Cu^{2+}} + 2\,\mathrm{Ag} & E^\circ = 0.80 - 0.34 = 0.46\ \mathrm{V} \end{array}$$

This treatment also allows determination of the direction of electron flow without the need for assigning signs to the two electrodes.

A powerful analogy can be drawn in lecture between water and electricity to help the students through the problems of signs. The amount of work done by water falling down a waterfall depends on the amount of water and the height through which the water falls (mgz). The amount of work done by an electric current depends on the amount of charge and the potential difference (nFE). Getting the students to think of cell potentials as an analogy to the height of a waterfall or as the distance between two points on a continuous scale minimizes sign mistakes. The analogy of charge to a quantity of water tends to minimize the mistake of wanting to use n in calculating cell potentials.

Teaching Tips

- Use the concept of half reactions to emphasize that oxidation cannot occur without reduction, and vice-versa.

- Students need help identifying the half reactions. Writing half reactions is dependent on correctly assigning oxidation states in compounds.

- Be sure that the students balance both atoms and charge in the half reactions. A common error is for students to place the electrons on the wrong side of the half reaction.

- Remind students to equalize the number of electrons transferred. They often forget to do this when they begin to work with these equations. Explain to them that the electrons are transferred and so must be accounted for. This means the number gained must equal the number lost.

- These reactions take time and practice to learn to balance. Students need to pay close attention to the details of the balancing process. Learning to use the method of half reactions on relatively easy equations makes it easier to balance more difficult equations later.

- Explain to students that electrochemical cells are different from ordinary oxidation-reduction reactions because the half reactions are separated from each other.

- Concentration cells can be explained by considering entropy. The greatest entropy exists when the two halves of the cell are mixed to the same concentration. Thus, having differing initial concentrations provides a driving force for the process.

- The lead storage battery is a practical example of electrolysis. Many students know that these batteries are “recharged” when the automobile engine runs. This is a good starting point for a discussion of electrolysis.

CHAPTER NINETEEN: THE NUCLEUS: A CHEMIST'S VIEW

Chapter Learning Goals:

Section One:	To relate the stability of a nucleus to the number of protons and neutrons. To classify the types of radioactive decay.
Section Two:	To define and show how to calculate the half-life of a radioactive nuclide.
Section Three:	To show how one element may be changed into another using particle bombardment.
Section Four:	To discuss radioactivity detection devices. To show how objects can be dated using radioactive decay.
Section Five:	To discuss the thermodynamic stability of the nucleus. To show how to calculate nuclear bonding energies.
Section Six:	To explore the energetics of nuclear fission and nuclear fusion. To describe how a nuclear reactor works.
Section Seven:	To show how radiation damages human tissue.

Chapter 19 is the only chapter in which the primary focus is on the nucleus. Alternative placement of the topics in this chapter were mentioned earlier: nuclear reactions in Chapter 2, rates of decay in Chapter 12.

There are some important applications of radioactivity. The biological sciences depend heavily on the use of radio-tracers. Analytical techniques, such as radioimmunoassay, use radioactive substances. Some of the more chemical applications of nuclear chemistry are covered in the chapter. The degree to which this chapter is covered will very much depend on the nature of the course being taught.

Teaching Tips

- Balancing nuclear equations provides an excellent opportunity to review balancing by atoms, charge, and the meanings of the mass number and atomic number in element notation.

- The students should be aware that in a nuclear decay the product is a different element. Therefore we cannot balance the equation by balancing the atoms.

- Remind students that in forming ions the number of electrons changes but the number of proton (and therefore the type of atom) remains the same. In nuclear equations if the number of protons is changed the atoms identity is also changed.

- Alchemists attempted to convert base metals (such as lead) to precious metals (such as gold) by chemical means. We now know that this can be accomplished only by nuclear reactions, which is not practical.

- Nuclear transformations are balanced in the same way as nuclear decay reactions.

- Radioisotopes are found in many household devices. Some examples are starters for fluorescent lamps and appliances (promethium-147, krypton-85 and thorium-232); antistatic devices (polonium-210); smoke detectors (americium-241).

- Carbon-14 is produced naturally in the environment. When it is incorporated into carbon dioxide it can be fixed as a carbohydrate in plants. Carbon dioxide containing carbon-14 is an example of a natural source of radioactivity in the environment with which we come in contact.

- The preservation of food by using gamma radiation was first approved by the Food and Drug Administration in 1963. The first foods approved for radiation preservation were wheat flour and potatoes. Not until 1986 could fruits and vegetables be preserved by radiation.

- Table 19.7 gives sources and amounts of radiation exposure for humans. This can be used in a discussion of the natural sources and exposure to radioactivity expected in our lives.

- The use of nuclear imaging techniques is a good starting point for a discussion of the hazards and benefits of nuclear medicine.

CHAPTER TWENTY: THE REPRESENTATIVE ELEMENTS

Chapter Learning Goals:

Section One:	To give an overview of the representative elements. To discuss the relationship of the first member to the rest of the group. To give a general introduction to the commercial preparation of some elements.
Section Two:	To consider the thermodynamics and rates of the reactions of alkali metals with water. To discuss the oxides of alkali metals.
Section Three:	To describe the production and uses of hydrogen. To discuss types of hydrides.
Section Four:	To describe the general chemical properties of the alkaline earth metals. To discuss three-center bonding in beryllium compounds.
Section Five:	To show the general trend from nonmetallic to metallic

	behavior in Group 3A.
Section Six:	To contrast the chemistry of carbon with that of silicon and the other members of Group 4A. To describe carbon oxides. To discuss the characteristics of the +2 and +4 oxidation states of tin and lead.
Section Seven:	To give an overview of the characteristics of the Group 5A elements, with emphasis on bonding.
Section Eight:	To discuss the kinetic and thermodynamic importance of the strength of the triple bond in the nitrogen molecule. To introduce the nitrogen cycle. To describe the structures and uses of major nitrogen-containing compounds.
Section Nine:	To characterize the forms of elemental phosphorus. To describe the preparation and structure of the major compounds of phosphorus.
Section Ten:	To discuss trends in the chemistry of the Group 6A elements.
Section Eleven:	To characterize the forms of elemental oxygen.
Section Twelve:	To characterize the elemental forms of sulfur. To describe the bonding and structure of the oxycompounds of sulfur.
Section Thirteen:	To describe the bonding in compounds of halogens with other nonmetals. To explain the acid strengths of the hydrogen halides. To discuss the chemistry of oxyhalogen compounds. To give the structures of interhalogen compounds.
Section Fourteen:	To describe the uses of the noble gases. To discuss the structures of xenon compounds.

Some examples of how topics from Chapter 20 can be integrated into some earlier chapters have already been discussed. The material in these chapters is organized by family in the periodic table. Many of the end-of-chapter exercises deal with the applications of chemical principles to these elements.

There are several other places in the text where topics from these chapters can be used. There is a short discussion of atmospheric chemistry at the end of Chapter 5, Gases. Instructors choosing to go into greater detail can use material in the chemistry of nitrogen, oxygen, and sulfur. Alternatively, atmospheric chemistry using information can be used to illustrate kinetics. Atmospheric chemistry is controlled to a much greater extent by kinetics than equilibrium. The catalytic destruction of the ozone layer by chlorine from chlorofluorocarbons is an important example that can easily get the students' attention.

Many of the elements that are important in high-technology materials such as composites, ceramics, photonic devices, semiconductors, and superconductors are representative elements. Instructors can use material from these chapters to illustrate applications of solids in Chapter 10. Because of the discussion of interparticle forces, Chapter 10 is a pivotal chapter in any general chemistry course and is a crossroads for moving into a variety of important applications.

Teaching Tip

- Use this chapter as a good summary of many of the concepts discussed in other chapters. For example, Section 20.1 requires the students to recall Section 7.12.

CHAPTER TWENTY-ONE: TRANSITION METALS AND COORDINATION CHEMISTRY

Chapter Learning Goals:

Section One:	To discuss electron configurations and general trends in the properties of the first-row transition metals. To introduce the 4d and 5d transition series.
Section Two:	To discuss some chemistry of the 3d transition metals.
Section Three:	To define terms used in coordination chemistry. To discuss some common chelates. To give rules for naming coordination compounds.
Section Four:	To introduce and illustrate types of structural isomerism and stereoisomerism. To relate molecular structure to optical activity.
Section Five:	To use the localized electron model to explain the interactions between a metal ion and ligands.
Section Six:	To describe the crystal field model and use it to explain the magnetism and colors of coordination complexes.
Section Seven:	To describe the function and structures of some biologically important coordination complexes.
Section Eight:	To describe processes for concentrating minerals in ores. To discuss the major processes in hydrometallurgy. To describe the blast furnace, open hearth, and basic oxygen processes for steelmaking. To discuss the thermodynamics of steelmaking.

We have previously mentioned that material from this chapter can be used in conjunction with a discussion of metallurgy (Chapter 10) or complex ions (Chapter 16). Another alternative is to cover transition metal chemistry after Chapter 9 to illustrate bonding models.

Two further points should be made concerning this chapter. First, the only discussion of electron configurations of transition metal ions appears in Section 21.1. Only the electron configurations of ions of representative elements are considered in the discussion of ionic bonding in Chapter 8. Second, the concept of isomerism is introduced in this chapter. Structural isomerism and stereoisomerism are illustrated with examples of transition metal complexes. Instructors who wish to introduce isomerism during earlier coverage of bonding may want to use Section 21.4. Alternatively, organic examples can be used for an early introduction to isomers. They are covered in Chapter 22 (structural and geometrical isomerism). Optical isomerism and chirality in organic compounds are introduced in Section 22.6 (carbohydrates).

Teaching Tips

- A compound such as $[Co(NH_3)_5Cl]Cl_2$ requires the students to have an understanding of polar covalent bonds (in the NH_3), coordinate covalent bonds (between the NH_3 and Cl^- ligands and the metal ion), and ionic bonding (between the chloride ions and the complex ion).

- It is helpful for many students to make physical models of the different types of isomers. This is especially true for the optical isomers with bidentate ligands such as shown in Figure 21.16.

- The crystal field model provides another opportunity to discuss the nature of science and the use and purpose of models in science. For example, we begin by defining ligands as negative point charges (even though this is not technically true). We use this definition when developing the model.

- Figure 21.21 shows a problem with the ligand as negative point charge definition. That is, all d-orbitals increase in energy. From discussion in Chapter 6 and Chapter 17, the students should realize that this should not occur. This shows us that the crystal field model, while correctly discussing splitting (and thus the color of such compounds), is a poor model for discussing energy.

- Consider using the color wheel to discuss complementary colors. Thus, red is complementary to green, orange is complementary to blue, and yellow is complementary to violet.

CHAPTER TWENTY-TWO: ORGANIC AND BIOLOGICAL MOLECULES

Chapter Learning Goals:

Section One:	To describe the nomenclature system for organic chemistry. To discuss isomerism in organic molecules. To give the rules for naming alkanes.
Section Two:	To introduce unsaturated hydrocarbons. To discuss isomerism in alkenes. To describe common reactions involving multiple carbon-carbon bonds.
Section Three:	To discuss the structures and reactivities of aromatic hydrocarbons.
Section Four:	To introduce the basic functional groups and give a few characteristics of each one.
Section Five:	To show how polymers are formed from monomers via two types of polymerization reactions.
Section Six:	To describe the levels of structure and the functions of proteins. To describe the structures and isomerism of simple carbohydrates. To describe the structures and functions of polymers of glucose. To describe the structures and functions of nucleic acids.

Chapter 22 emphasizes the structures of several classes of organic compounds. Instructors can use this chapter to lead to several applications. Energy and fuels can be emphasized by following up a discussion of hydrocarbons with petrochemicals (Section 22.4).

The structures and properties of synthetic polymers can be covered by emphasizing Section 22.5. Organic chemistry can be introduced after a discussion of bonding or, particularly if synthetic polymers are stressed, as an illustration of using interparticular forces in correlating structure and properties.

Optical isomerism in organic compounds and the chiral carbon atom are introduced in the section on carbohydrates (22.6). There are end-of-chapter problems on optical activity that deal only with simpler organic compounds for use by instructors who wish to introduce the topic to students without going into a detailed discussion of carbohydrates.

Some of the end-of-chapter exercises on nucleic acids deal with the genetic code and errors in the transmission of genetic information. The information science aspects of nucleic acids are an interesting application of entropy.

Teaching Tips

- To introduce the topic of organic chemistry and its impact on our lives ask the students what the classroom would look like without carbon containing compounds. You will need to assist them in figuring out what things contain carbon, but it will not be long before they can appreciate how many of the objects surrounding us are organic compounds.

- You will find a good set of three-dimensional models very helpful in presenting this material. Use the models to illustrate the shapes of the molecules discussed in this section. The tetrahedral shape of methane is difficult to visualize without a model. Inexpensive classroom sized models can be built from painted Styrofoam balls and wooden dowels.

- Strings of children’s pop beads can be used to illustrate that bending or twisting the string does not produce a change in the structure. To create a different order of beads you must take apart the chain and put it together differently.

- Organic nomenclature can seem overwhelming at first. It is not possible to memorize all the names and formulas for organic molecules. Students need to learn the names for the basic root and substituents. Once these are learned naming organic compounds becomes a problem-solving activity.

- Be sure to include some examples where the longest carbon chain is not horizontal. Students find it difficult to understand that a carbon chain with several bends in it is not branched. It is helpful to tell the students that the longest chain is the one which they can follow without lifting their pencil along the carbons. Circling or highlighting the longest chain of carbons can help when you are teaching the students to identify the substituents.

- It is helpful to show students how to think about isomer problems. An organized approach for drawing and naming isomers makes the process much simpler. Start with the longest carbon chain and remove one methyl group at a time and try to place it on as many new carbons as possible without producing duplicates. Naming each molecule will help to eliminate duplicates. If the names are the same then the molecules are the same even if they look different in the drawing.

- In substitutions reactions be sure to remind the students that any one of the hydrogens can be substituted. Not all substitutions will produce different products. This can be confusing to some students. Use three-dimensional models to illustrate this one molecule at a time.

- A model that shows the difference between cyclohexane (a puckered ring) and benzene (a flat ring) would be useful in helping students understand aromatic hydrocarbons.

- You may wish to discuss the delocalization of the electrons in the benzene ring as an example of how our bonding model is limited. Drawing the benzene ring structure with the circle inside the hexagon reminds us that benzene behaves differently than a carbon molecule with a double bond.

- When determining whether an alcohol is primary, secondary, or tertiary the focus should be on the carbon where the –OH is attached. Students often focus on the –OH group itself instead and this does not tell anything about the classification.

- Section 22.6 is an excellent place to review what causes a side group to be polar or nonpolar. Refer the students back to the ideas they learned in bonding and work with them to determine why each of the side chains are classified as polar or nonpolar.

- The term "hydrophobic" is actually a misnomer. While the non-polar side chains do not mix well with water, they do not repel the water. Water molecules are more attracted to other water molecules than to the non-polar side chains.

- An excellent model of secondary structure is the coiled cord from a telephone. Another possibility is to use the children's pop beads and twist them into a helix to represent secondary structure while showing the primary structure by the order of the pop beads. A folded piece of paper can represent the pleated sheet secondary structure.

- You can use the analogy of a ladder to help students understand the structure of DNA. The sides of the ladder represent the alternating sugar and phosphate groups and the rungs of the ladder represent the base pairs that are hydrogen bonded together.

- Fats containing more unsaturated fatty acids are liquid at room temperature.

- Lipids are different from most other organic and biomolecules because they are distinguished not by a particular functional group but by their solubility in nonpolar solvents.

EXPERIMENT 2

The Use of Volumetric Glassware

A. General Notes

The correct measurement of volume is essential to students' future performance in the general chemistry laboratory, and this experiment will give the instructor a chance to determine which students may have difficulties in the future. In particular, students will need a great deal of help in learning to read measuring scales correctly. It is especially difficult for students to appreciate the concept of estimating between the smallest scale divisions of a device to obtain the last figure of the measurement.

As this experiment progresses, the instructor should examine student data *as they are being recorded* by circulating in the lab. In this manner, if it is discovered that a student is reading or recording data incorrectly, the student can be corrected *before* he or she mentally reinforces the incorrect technique. For example, if the instructor notices that a student is recording buret volumes to only *one* decimal place, it is crucial to explain to the student *on the spot*, while the buret is in front of him or her, how and why we record such volume readings to a second, estimated decimal place.

This experiment introduces students to volumetric measuring glassware of both low and high precision. In particular, the use of a volumetric transfer pipet is covered since a number of the later experiments make use of such pipets. Because several types of pipet safety bulbs may be available, the instructor should demonstrate the type of bulb the students will actually use.

Estimated time to complete: 75–90 minutes

Since students may not have yet covered the concept of density in their lecture course, it is suggested that the calculations of pipet volumes/mass of water transferred be covered at the chalkboard.

B. Materials Required (Given per 25 students)

Set-up of several different-sized graduated cylinders containing different amounts of colored water. The cylinders should be marked with a code letter for students to distinguish them. Either the cylinders should be sealed with a rubber stopper or plastic laboratory film to prevent evaporation, or the instructor should read and record the volume of liquid in the cylinders at the start of the lab period.

25 25-mL pipets

25 pipet safety bulbs

25 50-mL burets and buret stands/clamps

5–6 buret brushes

C. Answers to Pre-Laboratory Questions

1. Volumetric glassware must be scrupulously clean in order that the volumes indicated by the manufacturer will be correct. Washing such glassware after use will make it easier to clean for the next use, and also will prevent any corrosive materials from reacting with the glass of the container and changing the volume of the container. For example, sodium hydroxide reacts slowly with glass: if a pipet or buret was used to dispense a sodium hydroxide solution, not rinsing the pipet or buret to remove the sodium hydroxide would gradually change the volume of the pipet or buret.

2. This notation indicates that the pipet is calibrated to deliver 10.00 mL at 25°C.

3. Any air bubble in the tip of a buret must be removed before the buret is used for dispensing the liquid in the buret. If the air bubble were to be displaced during the dispensing of the liquid, this would introduce an error equal to the volume of the air bubble.

   For the 20-mL sample: [(0.5)/(20)](100) = 2.5% error

4. For the 1-mL pipet: [(5)(0.05)/(1.00)](100) = 25% error

   For the 5-mL pipet: [(5)(0.05)/(5.00)](100) = 5% error

   For the 10-mL pipet: [(5)(0.05)/(10.00)](100) = 2.5% error

5. A rubber safety bulb is always used to prevent aspiration of toxic or corrosive substances into the mouth.

D. Answers to Post-Laboratory Question

1. It is likely that students will conclude that the buret and the pipet permit similar precision, whereas the graduated cylinder permits less precision. Typically, a graduated cylinder is used when only an approximate volume is needed. A volumetric transfer pipet is used when a precise, specific volume is needed, and a buret may be used when a precise, nonspecific volume is required.

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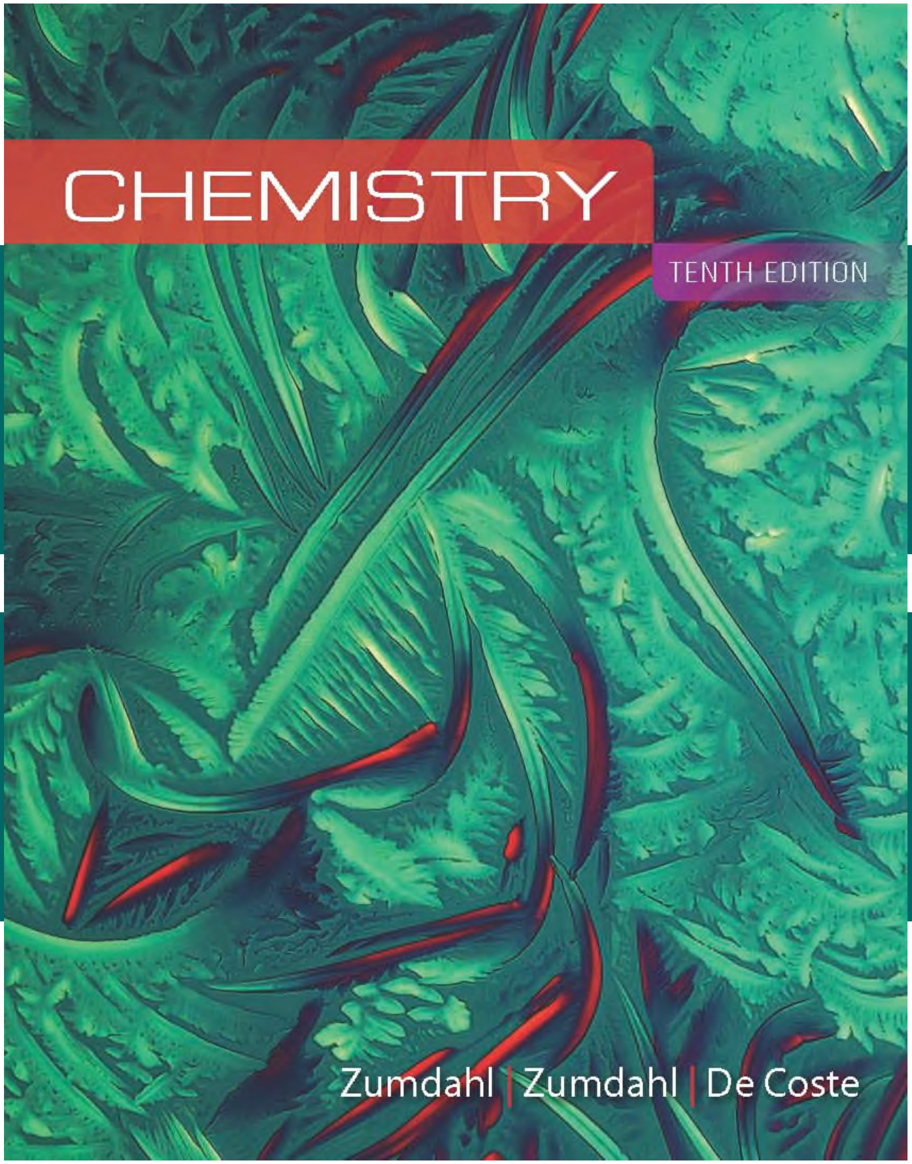


Chapter 2

Atoms, Molecules, and Ions

Chapter 2
Table of Contents

Section 2.1

The Early History of Chemistry

Early History of Chemistry

- Applications of chemistry before 1000 B.C.
  - Use of embalming fluids
  - Processing of natural ores to produce metals for ornaments and weapons
- The Greeks (400 B.C.)
  - Proposed that matter was composed of four fundamental substances (earth, fire, air, and water)

Early History of Chemistry (continued)

    - Considered the question of whether matter is infinitely divisible or is composed of small, indivisible particles
- Alchemists dominated the field of chemistry for the next 2000 years
    - Helped discover several elements
    - Learned to prepare mineral acids

Section 2.1

The Early History of Chemistry

Modern Chemistry

- Foundations were laid in the 16th century by:
  - Georg Bauer, who developed systematic metallurgy
  - Paracelsus, who discovered the medicinal application of minerals
- Robert Boyle (1627–1691)
  - Performed quantitative experiments to measure the relationship between the pressure and volume of air

Modern Chemistry (continued 1)

- Developed the first experimental definition of an element
    - A substance is an element unless it can be broken down into two or more simpler substances
- Held on to certain alchemists’ views
    - Metals were not true elements, and eventually, a method to change one metal to another will be found

Modern Chemistry (continued 2)

- 17th and 18th century
    - Interest in the phenomenon of combustion rose
    - Joseph Priestley discovered oxygen gas
    - Georg Stahl
        - Suggested that a substance called phlogiston flowed out of burning material
        - Postulated that substances that burn in closed containers eventually stop burning since the air in the container is saturated with phlogiston

Section 2.1

The Early History of Chemistry

Modern Chemistry (continued 3)

- Oxygen vigorously supported combustion and was thus supposed to be low in phlogiston
    - Was originally called dephlogisticated air

Antoine Lavoisier

- Explained the true nature of combustion
- Regarded measurement as the essential operation of chemistry
- Verified the law of conservation of mass
    - **Law of conservation of mass**: Mass is neither created nor destroyed in a chemical reaction
- Conducted quantitative experiments that showed that combustion involved oxygen, not phlogiston

Section 2.2
Fundamental Chemical Laws

Antoine Lavoisier (continued)

- Discovered that life was supported by a process that also involved oxygen and was similar in many ways to combustion

Section 2.2

Fundamental Chemical Laws

Joseph Proust

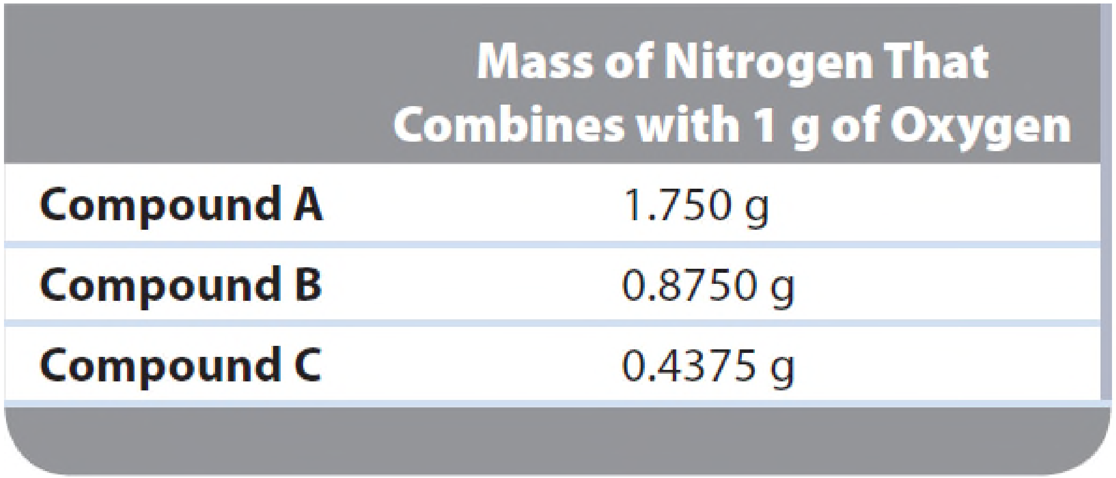
- Proposed the principle of the constant composition of compounds or Proust’s law or the law of definite proportion
  - **Law of definite proportion**: A given compound always contains exactly the same proportion of elements by mass

John Dalton's Law of Multiple Proportions

- When two elements form a series of compounds, the ratios of the masses of the second element that combine with 1 g of the first element can always be reduced to small whole numbers

Example 2.1 - Illustrating the Law of Multiple Proportions

- The following data were collected for several compounds of nitrogen and oxygen:



	Mass of Nitrogen That Combines with 1 g of Oxygen
Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

  - Show how these data illustrate the law of multiple proportions

Example 2.1 - Solution

- For the law of multiple proportions to hold, the ratios of the masses of nitrogen combining with 1 g of oxygen in each pair of compounds should be small whole numbers
  - Therefore, compute the ratios as follows:

$$\frac{A}{B} = \frac{1.750}{0.8750} = \frac{2}{1}$$

Example 2.1 - Solution (continued)

$$\frac{B}{C} = \frac{0.8750}{0.4375} = \frac{2}{1}$$

$$\frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

- These results support the law of multiple proportions

Join In (1)

- According to the law of definite proportions:
  a. If the same two elements form two different compounds, they do so in the same ratio
  b. It is not possible for the same two elements to form more than one compound
  c. The ratio of the masses of the elements in a compound is always the same
  d. The total mass after a chemical change is the same as before the change

Dalton’s Atomic Theory

- Each element is made up of tiny particles called atoms
- Atoms of a given element are identical
    - Atoms of different elements are different in some fundamental way or ways

Dalton's Atomic Theory (continued)

- Chemical compounds are formed when atoms of different elements combine with each other
  - A given compound always has the same relative numbers and types of atoms
- Chemical reactions involve reorganization of the atoms
  - Atoms themselves are not changed in a chemical reaction

Section 2.3
Dalton’s Atomic Theory

Table of Atomic Masses

- Dalton prepared the first table of **atomic masses** (**atomic weights**) based on the assumption that nature would be as simple as possible
  - Many masses were proved to be wrong because of Dalton’s incorrect assumptions about the formulas of certain compounds

Joseph Gay-Lussac

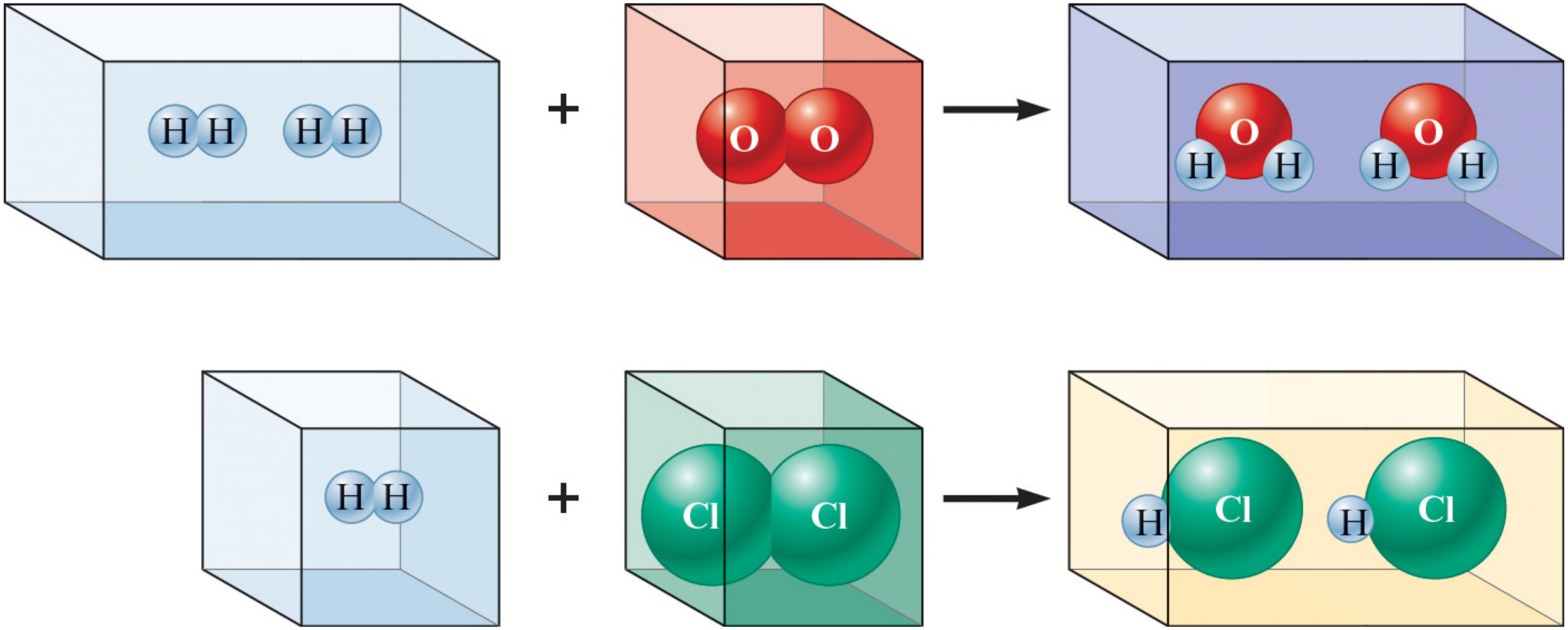
- Measured the volumes of gases that reacted with each other under the same temperature and pressure conditions

Avogadro's Hypothesis

- At the same temperature and pressure, equal volumes of different gases contain the same number of particles
    - Makes sense if the distances between the particles in a gas are very great compared with the sizes of the particles
    - Volume of a gas is determined by the number of molecules present, not by the size of the individual particles

Section 2.3
Dalton's Atomic Theory

Figure 2.5 - Representation of Combining Gases at the Molecular Level



The spheres represent atoms in the molecules

Section 2.3

Dalton’s Atomic Theory

Join In (2)

- How many of the following did Dalton not discuss in his atomic theory?
  - I. Isotopes
  - II. Ions
  - III. Protons
  - IV. Neutrons
  - V. Electrons

a. I

b. II

c. III

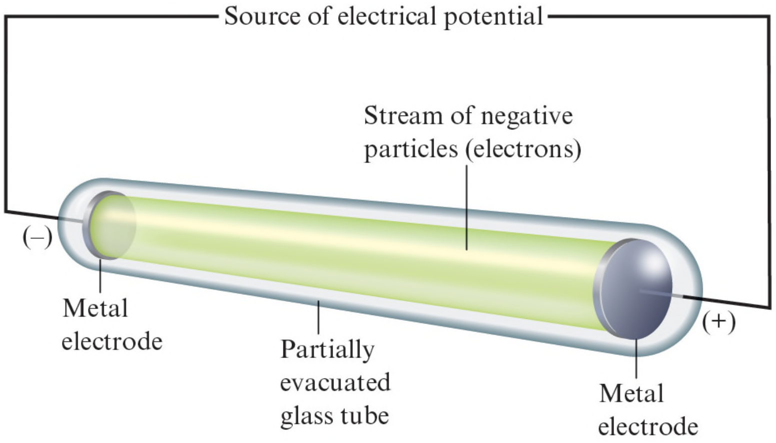
d. IV

e. V

Section 2.4
Early Experiments to Characterize the Atom

J. J. Thomson

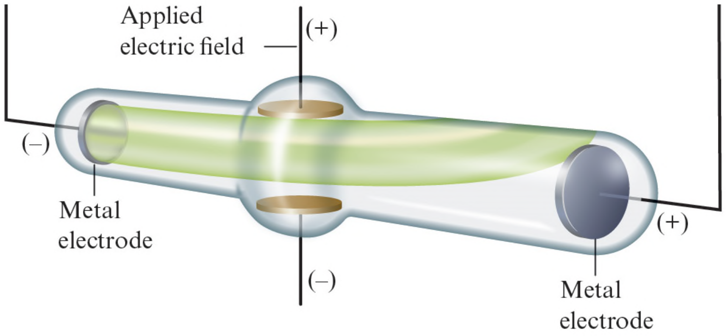
- Studied electric discharges in partially evacuated tubes called **cathode-ray tubes**



Section 2.4
Early Experiments to Characterize the Atom

J. J. Thomson (continued 1)

- Postulated that the cathode ray was a stream of negatively charged particles (**electrons**)
  - Cathode ray was produced at the negative electrode when high voltage was applied to the tube
    - Repelled by the negative pole of an applied electric field



J. J. Thomson (continued 2)

- Determined the charge-to-mass ratio of an electron

$$\frac{e}{m} = -1.76 \times 10^{8}\ \text{C/g}$$

  - *e* - Charge on the electron (in coulombs)
  - *m* - Electron mass (in grams)

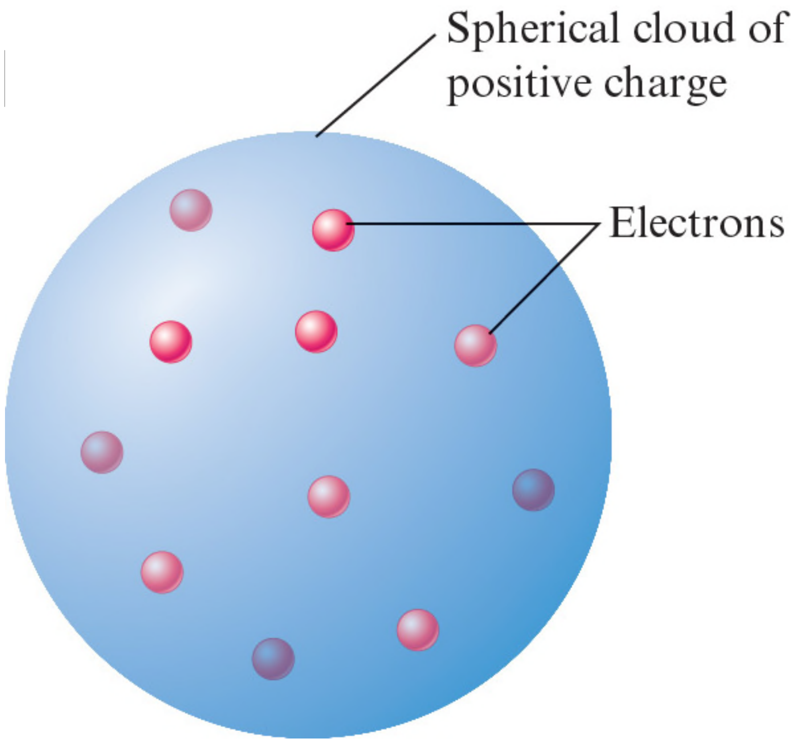
J. J. Thomson - Development of the Plum Pudding Model

- Primary goal of the cathode-ray tube experiments was to understand the structure of an atom
- Assumptions
  - All atoms must contain electrons
    - Electrons can be produced from electrodes made of various metals
  - Atoms must contain some amount of positive charge
    - Atoms were known to be electrically neutral

Section 2.4
Early Experiments to Characterize the Atom

J. J. Thomson’s Plum Pudding Model

- Atoms consist of a diffuse cloud of positive charge wit the negative electrons embedded randomly in it



Section 2.4
Early Experiments to Characterize the Atom

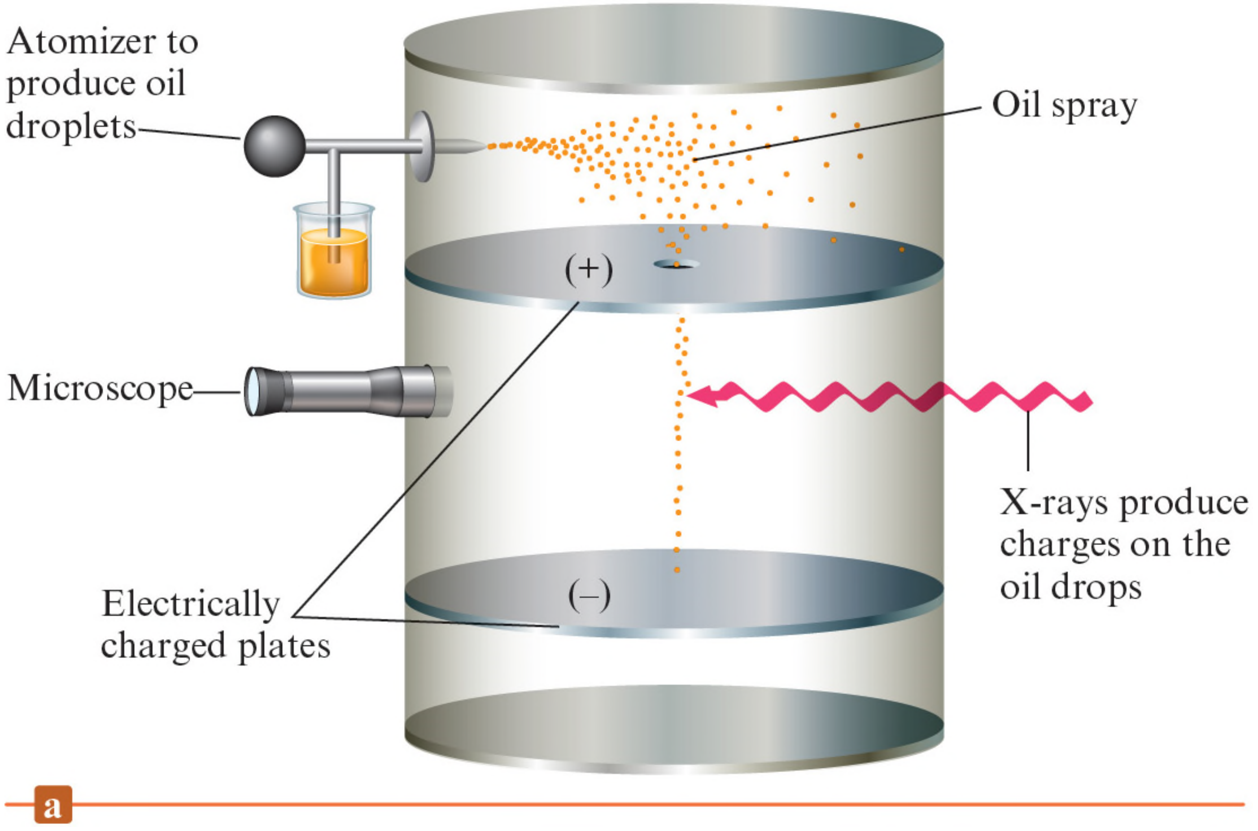
Robert Millikan

- Performed experiments involving charged oil drops, which helped determine the magnitude of electron charge
  - Used this value and the charge-to-mass ratio to calculate the mass of an electron as 9.11×10^{-31} kg

Section 2.4

Early Experiments to Characterize the Atom

Figure 2.10 (a) - Schematic Representation of the Apparatus Used to Determine the Charge on the Electron



Section 2.4

Early Experiments to Characterize the Atom

Henri Becquerel

- Discovered **radioactivity** by observing the spontaneous emission of radiation by uranium
  - Observed that a mineral containing uranium produces its image on a photographic plate in the absence of light

Section 2.4

Early Experiments to Characterize the Atom

Types of Radioactive Emission

- Gamma rays (γ)
  - High-energy light
- Beta particles (β)
  - High-speed electrons
- Alpha particles (α)
  - Particles with 2+ charge
  - Mass is 7300 times that of the electron

Rutherford’s Experiment

- Carried out to test the accuracy of Thomson’s plum pudding model
- Involved directing α particles at a thin sheet of metal foil
- Expectation
  - α particles will pass through the foil with minor deflections in their paths

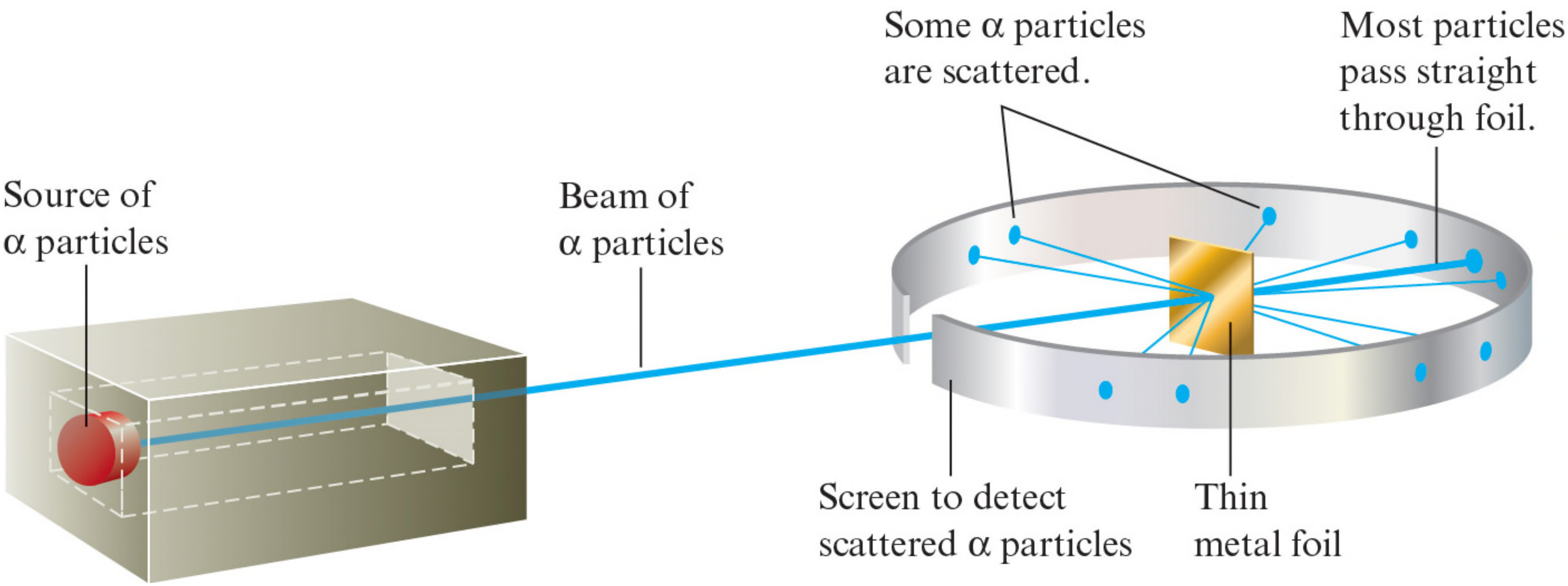
Rutherford’s Experiment - Results

- Most α particles passed through the foil
    - Atom is mostly open space
- Many particles were deflected at large angles
    - Including those that had a close encounter with the massive positive center of the atom
- Some particles were reflected
    - Including those that made a direct hit on the massive positive center

Section 2.4

Early Experiments to Characterize the Atom

Figure 2.12 - Rutherford's Experiment on α-Particle Bombardment of a Metal Foil



Section 2.4

Early Experiments to Characterize the Atom

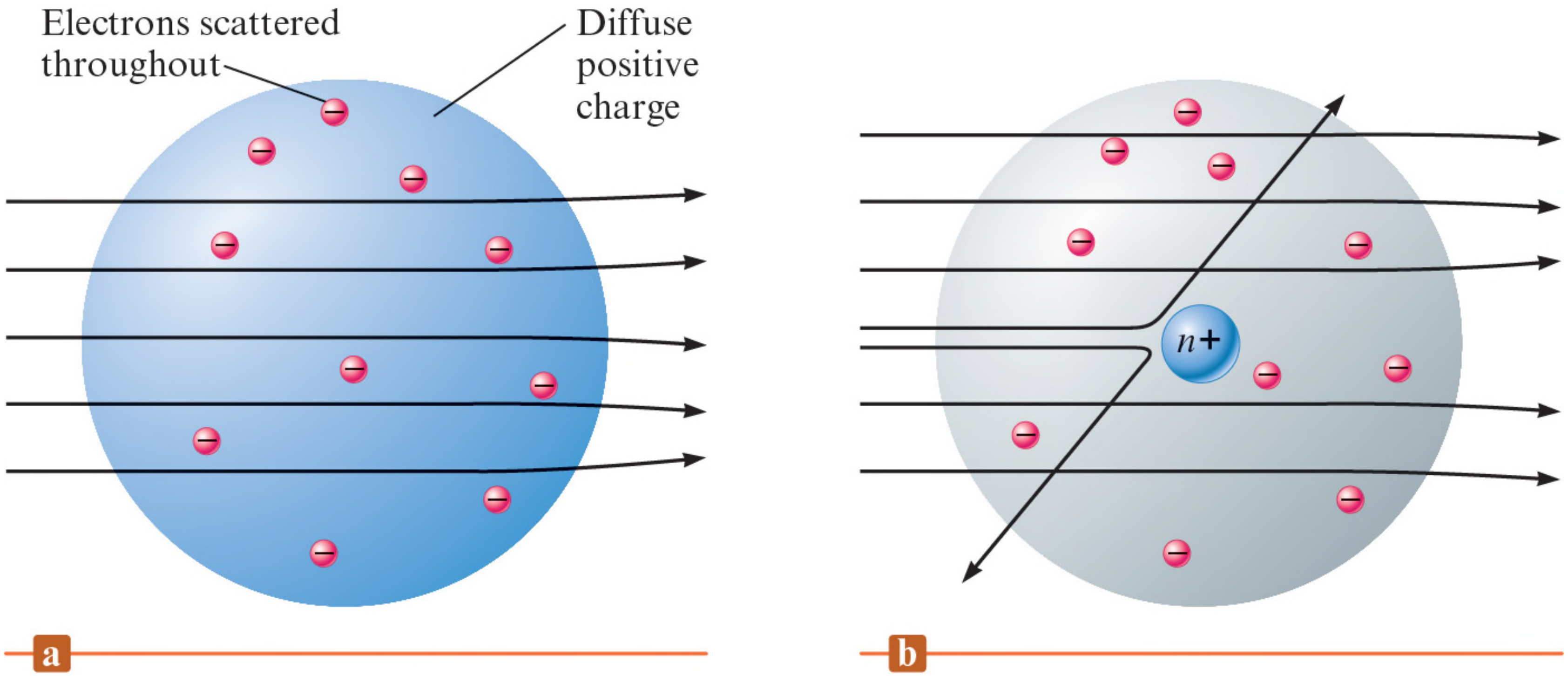
Rutherford's Experiment - Conclusions

- Result of the experiment can be explained in terms of a **nuclear atom**
  - Has a dense center of positive charge called the **nucleus** with electrons moving around the nucleus at a distance that is large relative to the nuclear radius

Section 2.4

Early Experiments to Characterize the Atom

Figure 2.13 - Rutherford's Experiment



The expected results of the metal foil experiment if Thomson's model were correct.

Actual results.

Section 2.4

Early Experiments to Characterize the Atom

Critical Thinking

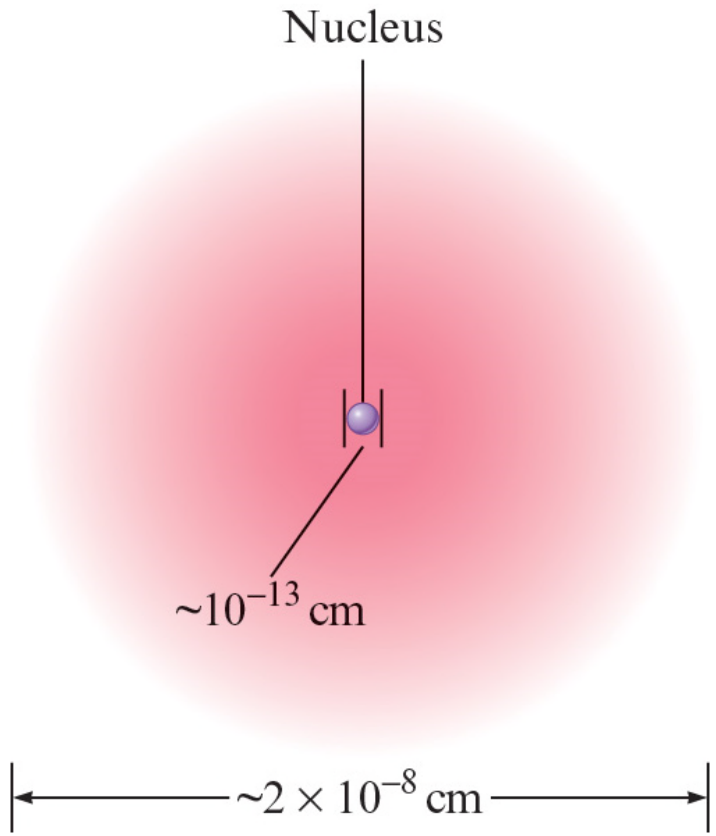
- You have learned about three different models of the atom: Dalton’s model, Thomson’s model, and Rutherford’s model
    - What if Dalton was correct? What would Rutherford have expected from his experiments with gold foil?
    - What if Thomson was correct? What would Rutherford have expected from his experiments with gold foil?

Atomic Structure

- Nucleus is assumed to contain:
    - **Protons**: Have a positive charge that is equal in magnitude to the electron’s negative charge
    - **Neutrons**: Have virtually the same mass as a proton but no charge
- Atoms of different elements, which have different numbers of protons and electrons, exhibit different chemical behavior

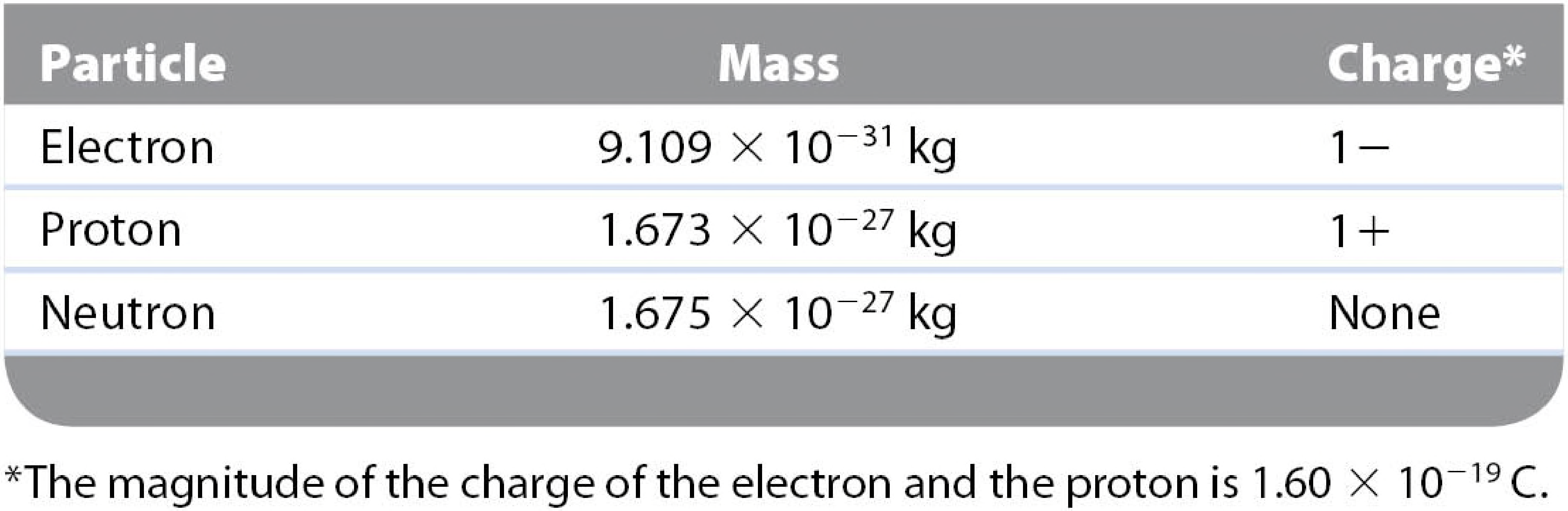
Features of the Nucleus

- Small in size compared to the overall size of the atom
- High density



Section 2.5
The Modern View of Atomic Structure: An Introduction

Table 2.1 - The Mass and Charge of the Electron, Proton, and Neutron



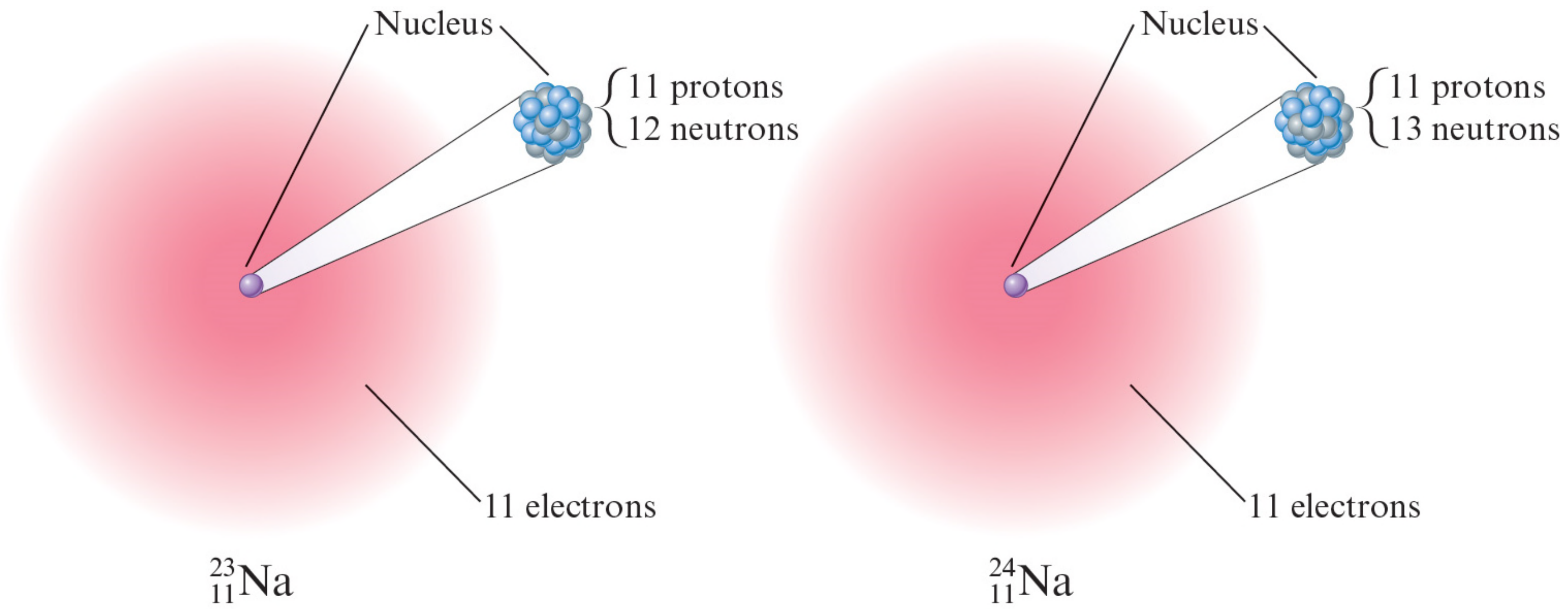
Particle	Mass	Charge*
Electron	9.109×10^{-31} kg	1−
Proton	1.673×10^{-27} kg	1+
Neutron	1.675×10^{-27} kg	None

*The magnitude of the charge of the electron and the proton is 1.60×10^{-19} C.

Isotopes

- Atoms with the same number of protons but different numbers of neutrons
- Depict almost identical chemical properties
- In nature, most elements contain mixtures of isotopes

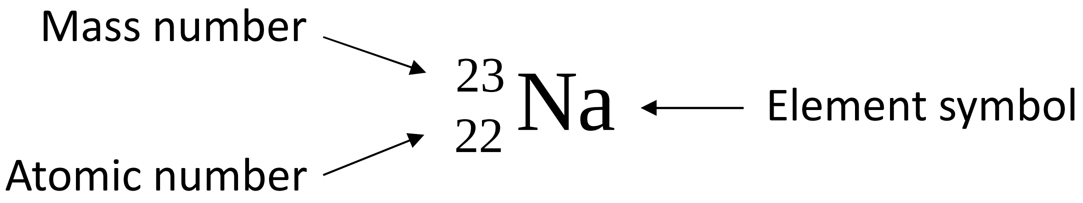
Figure 2.15 - Two Isotopes of Sodium



Section 2.5
The Modern View of Atomic Structure: An Introduction

Identifying Isotopes

- **Atomic number** (*Z*): Number of protons
    - Written as a subscript
- **Mass number** (*A*): Total number of protons and neutrons
    - Written as a superscript



$^{23}_{22}\mathrm{Na}$

Critical Thinking

- The average diameter of an atom is 2×10^{-10} m
  - What if the average diameter of an atom were 1 cm?
  - How tall would you be?

Interactive Example 2.2 - Writing the Symbols for Atoms

- Write the symbol for the atom that has an atomic number of 9 and a mass number of 19
  - How many electrons and how many neutrons does this atom have?

Interactive Example 2.2 - Solution

- The atomic number 9 means the atom has 9 protons
  - This element is called fluorine, symbolized by F
  - The atom is represented as follows:

$$^{19}_{9}\text{F}$$

    - The atom is called fluorine nineteen

Interactive Example 2.2 - Solution (continued)

- Since the atom has 9 protons, it also must have 9 electrons to achieve electrical neutrality
    - The mass number gives the total number of protons and neutrons, which means that this atom has 10 neutrons

Exercise

- How many protons and neutrons are in the nucleus of each of the following atoms?
  - In a neutral atom of each element, how many electrons are present?

1.	^{79}Br	**35 p, 44 n, 35 e**
2.	^{81}Br	**35 p, 46 n, 35 e**
3.	^{239}Pu	**94 p, 145 n, 94 e**
4.	^{133}Cs	**55 p, 78 n, 55 e**

Join In (3)

- The element rhenium (Re) exists as 2 stable isotopes and 18 unstable isotopes
  - The nucleus of rhenium-185 contains:
    a. 75 protons and 75 neutrons
    b. 75 protons and 130 neutrons
    c. 130 protons and 75 neutrons
    d. 75 protons and 110 neutrons

Join In (4)

- Which of the following statements are true?
  I. The number of protons is the same for all neutral atoms of an element
  II. The number of electrons is the same for all neutral atoms of an element
  III. The number of neutrons is the same for all neutral atoms of an element

Section 2.5

The Modern View of Atomic Structure: An Introduction

Join In (4) (continued)

a. I, II, and III are true
b. Only I and II are true
c. Only II and III are true
d. Only I and III are true
e. I, II, and III are false

Chemical Bonds

- Forces that hold atoms together in compounds
- **Covalent bond**: Formed by sharing electrons
  - Resulting collection of atoms is called a **molecule**, which can be represented in the following ways:
    - Chemical formula
    - Structural formula
    - Space-filling model
    - Ball-and-stick model

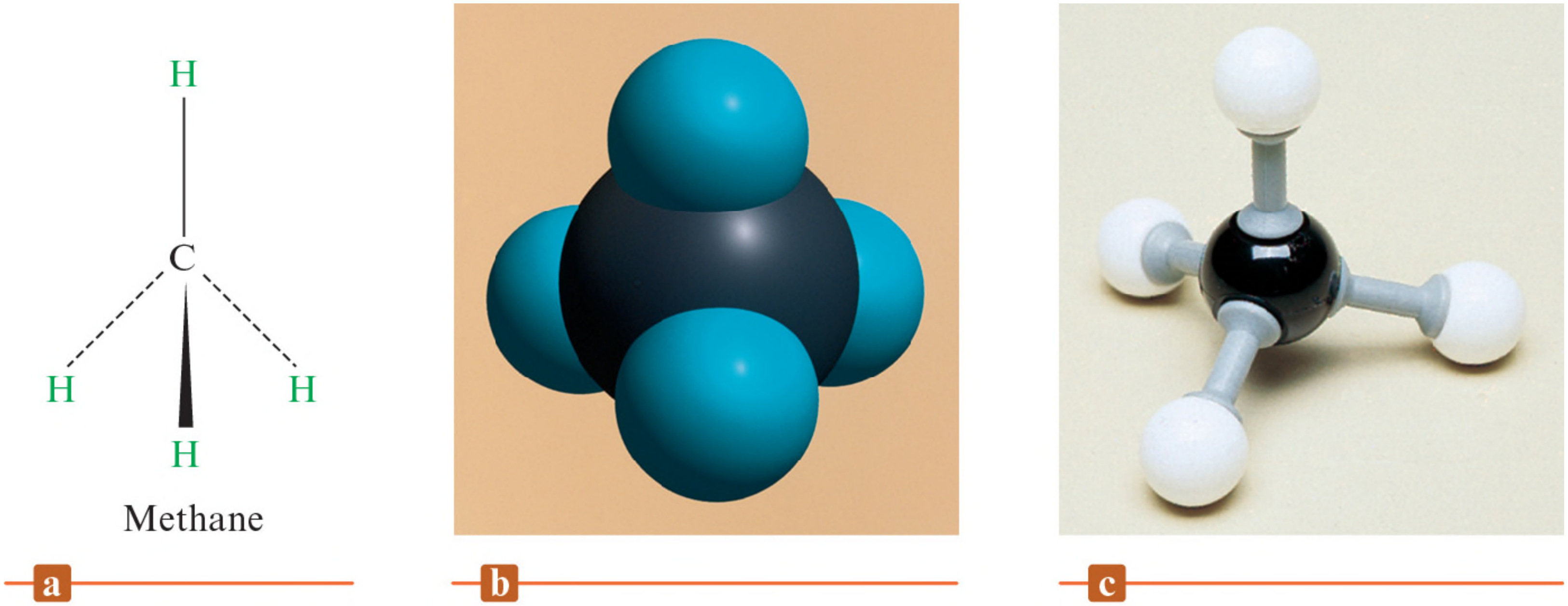
Chemical Formula

- Uses symbols of elements to indicate types of atoms present in the molecule
- Subscripts indicate the relative number of atoms
- Example
  - Formula for carbon dioxide is CO_2
    - Implies that each molecule of CO_2 contains one atom of carbon and two atoms of oxygen

Other Methods for Representing a Molecule

- **Structural formula**
    - Depicts individual bonds in a molecule
    - May or may not indicate the actual shape of the molecule
- **Space-filling model**
    - Illustrates the relative sizes of atoms and their relative orientation in the molecule
- **Ball-and-stick model**

Figure 2.16 - Molecular Representations of Methane



Structural formula

Space-filling model

Ball-and-stick model

Section 2.6
Molecules and Ions

Ion

- Atom or group of atoms with a net positive or negative charge
  - **Cation**: Positive ion formed by losing electrons
  - **Anion**: Negative ion formed by gaining electrons
- **Ionic bonding**: Force of attraction between oppositely charged ions

Ion Formation - Example

- Sodium chloride
  - Formed when a neutral Cl and Na react
  - Electron is transferred from a Na atom to a Cl atom

$$\text{Cation formation: Na} \longrightarrow \text{Na}^{+} + \text{e}^{-}$$

$$\text{Anion formation: Cl} + \text{e}^{-} \longrightarrow \text{Cl}^{-}$$

Ionic Solids

- Solids containing oppositely charged ions
- Can consist of:
  - Simple ions
    - Example - Sodium chloride (NaCl)
  - **Polyatomic ions**: Contain many atoms
    - Example - Ammonium nitrate (NH_4NO_3) contains ammonium ions (NH_4^+) and nitrate ions (NO_3^-)

Exercise

- Would you expect each of the following atoms to gain or lose electrons when forming ions?
  - What ion is the most likely in each case?

a.	Ra	**Loses 2 e^- to form Ra^{2+}**
b.	In	**Loses 3 e^- to form In^{3+}**
c.	P	**Gains 3 e^- to form P^{3-}**
d.	Te	**Gains 2 e^- to form Te^{2-}**

Join In (5)

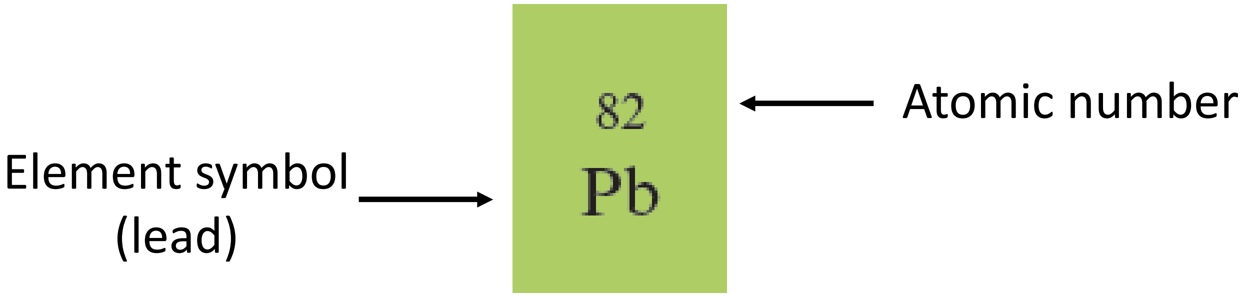
- How is an ion formed?
  - a. By either adding or subtracting protons from the atom
  - b. By either adding or subtracting electrons from the atom
  - c. By either adding or subtracting neutrons from the atom
  - d. All of these are true
  - e. Two of these are true

Join In (6)

- An element’s most stable ion forms an ionic compound with chlorine and has the formula XCl_2
  - If the ion has 36 electrons, what is the element from which the ion comes?
    a. Kr
    b. Se
    c. Sr
    d. Rb
    e. None of these

The Periodic Table

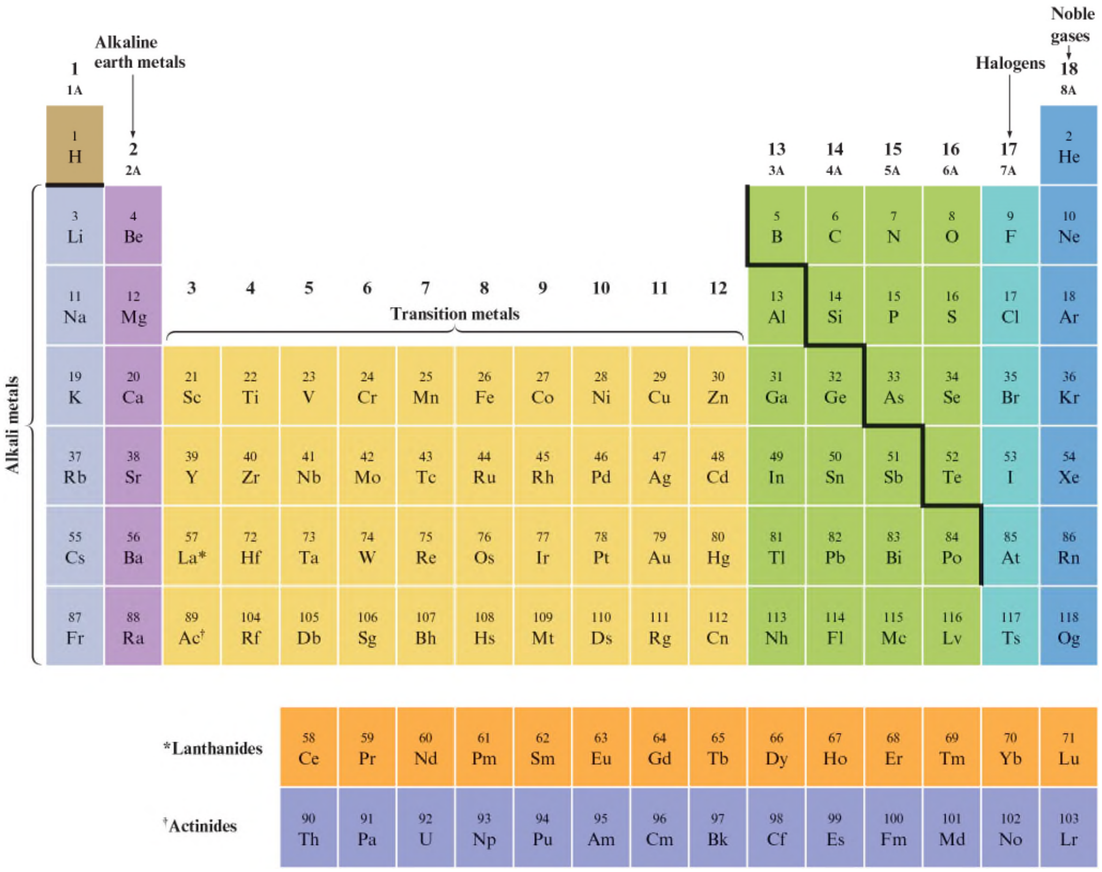
- Provides information about all known elements
    - Letters in boxes are symbols of elements
    - Number above every symbol is the element’s atomic number



Section 2.7

An Introduction to the Periodic Table

Figure 2.19 - The Periodic Table



1 1A	2 2A	3	4	5	6	7	8	9	10	11	12	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

*Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Metals and Nonmetals

Metals

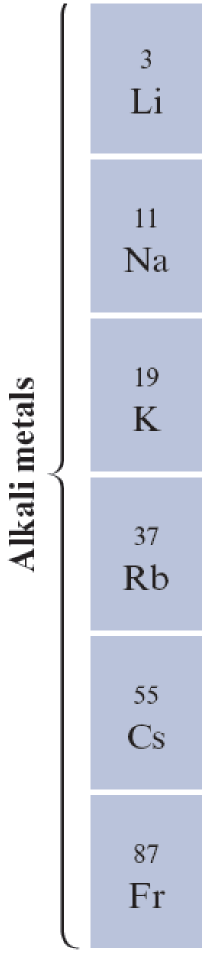
- Efficient conductors of heat and electricity, malleable, and ductile
- Have a lustrous appearance
- Tend to lose electrons to form positive ions

Nonmetals

- Lack the physical properties that characterize metals
- Tend to gain electrons in reactions with metals to form negative ions
- Often bond to each other by forming covalent bonds

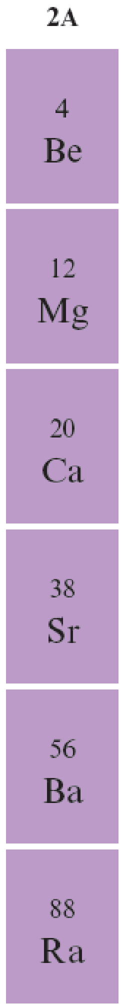
Structure of the Periodic Table: Groups or Families

- Elements in the vertical columns with similar chemical properties
- **Alkali metals**
  - Members of Group 1A
  - Very active elements that readily form ions with a 1+ charge when they react with nonmetals



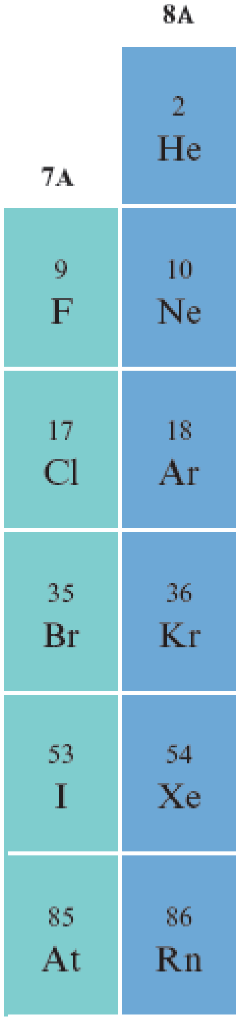
Structure of the Periodic Table: Groups or Families
(continued 1)

- **Alkaline earth metals**
  - Members of Group 2A
  - Form ions with a 2+ charge when they react with nonmetals



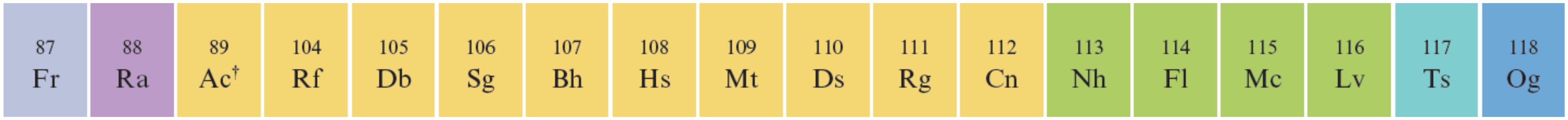
Structure of the Periodic Table: Groups or Families
(continued 2)

- **Halogens**: Members of Group 7A
  - Form diatomic molecules
  - React with metals to form salts containing ions with a 1– charge (exception - Astatine)
- **Noble gases**: Members of Group 8A
  - Exist under normal conditions as monatomic gases
  - Have little chemical reactivity



Structure of the Periodic Table: Periods

- Horizontal rows of elements
  - Horizontal row 1 is called the first period, row 2 is called the second period, and so on



87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

Join In (7)

- _____ form ions with a 2+ charge when they react with nonmetals
  a. Alkali metals
  b. Alkaline earth metals
  c. Halogens
  d. Noble gases
  e. None of these

Section 2.8
Naming Simple Compounds

Binary Compounds

- Composed of two elements
- Include covalent and ionic compounds
  - **Binary ionic compounds**: Contain a cation, which is written first in the formula, and an anion

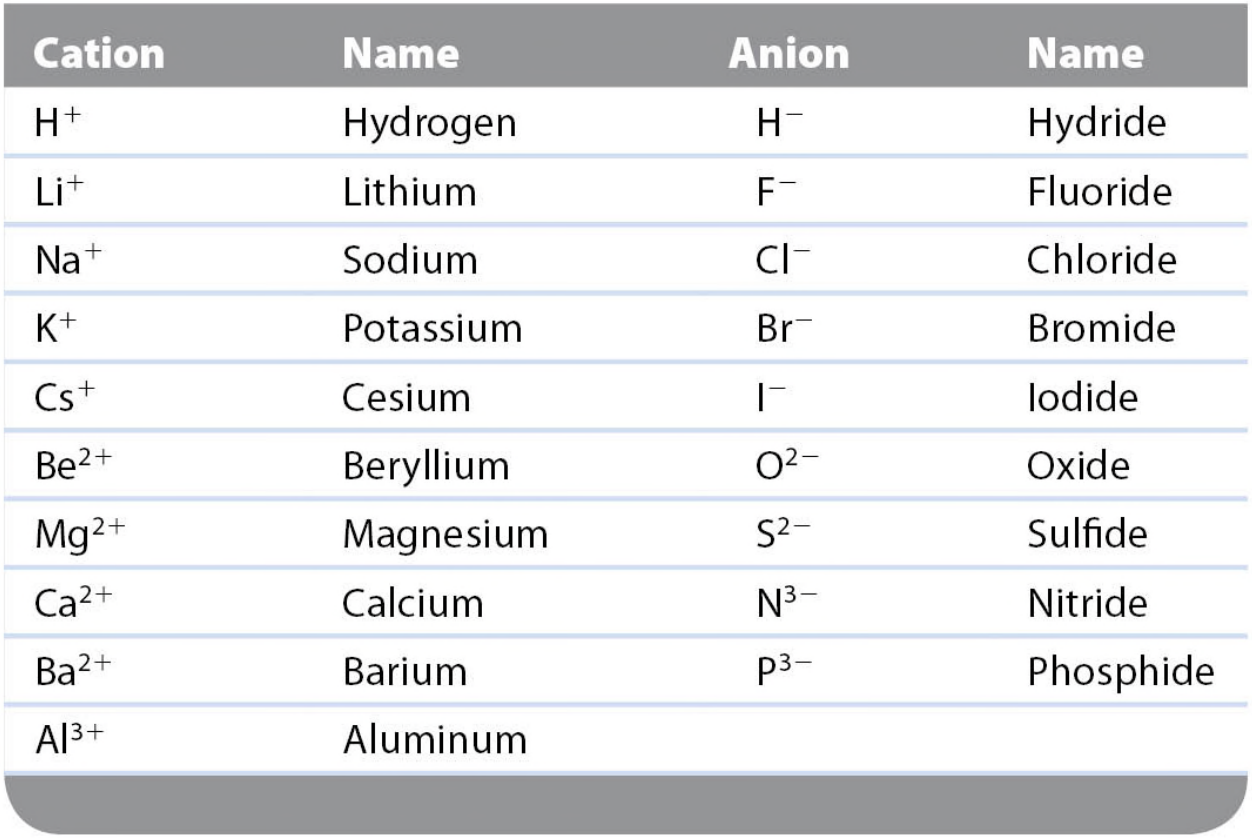
Naming Binary Ionic Compounds (Type I)

- Cation is always named first and the anion second
- Monatomic cation takes its name from the name of the parent element
- Monatomic anion is named by taking the root of the element name and adding -ide

Section 2.8

Naming Simple Compounds

Table 2.3 - Common Monatomic Cations and Anions



Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Cesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulfide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminum		

Interactive Example 2.3 - Naming Type I Binary Compounds

- Name each binary compound
  1. a. CsF
  2. b. $AlCl_3$
  3. c. LiH

Section 2.8

Naming Simple Compounds

Interactive Example 2.3 - Solution

a. CsF is cesium fluoride
b. $AlCl_3$ is aluminum chloride
c. LiH is lithium hydride

- Notice that, in each case, the cation is named first and then the anion is named

Binary Ionic Compounds (Type II)

- Nomenclature for metals that form more than one type of cation
  - Charge of the metal cation is indicated by a Roman numeral
  - Common metals that don’t require Roman numerals
    - Group 1A elements, which form only 1+ ions
    - Group 2A elements, which form only 2+ ions
    - Aluminum, which forms only Al^{3+}
    - Silver and zinc (Ag^{+} and Zn^{2+})

Section 2.8
Naming Simple Compounds

Binary Ionic Compounds (Type II) (continued)

- Alternative nomenclature
    - Ion with the higher charge has a name ending in -ic, and the one with the lower charge has a name ending in -ous

Interactive Example 2.5 - Naming Type II Binary Compounds

1. Give the systematic name for each of the following compounds:
   a. CuCl   b. HgO   c. Fe_2O_3
2. Given the following systematic names, write the formula for each compound:
   a. Manganese(IV) oxide
   b. Lead(II) chloride

Interactive Example 2.5 - Solution

- All of these compounds include a metal that can form more than one type of cation
  - Thus we must first determine the charge on each cation
    - This can be done by recognizing that a compound must be electrically neutral; that is, the positive and negative charges must exactly balance

Interactive Example 2.5 - Solution (1)

1.	**Formula**	**Name**	**Comments**
a.	CuCl	Copper(I) chloride	Because the anion is Cl^-, the cation must be Cu^+ (for charge balance), which requires a Roman numeral I
b.	HgO	Mercury(II) oxide	Because the anion is O^{2-}, the cation must be Hg^{2+} [mercury(II)]
c.	Fe_2O_3	Iron(III) oxide	The three O^{2-} ions carry a total charge of 6–, so two Fe^{3+} ions [iron(III)] are needed to give a 6+ charge

Interactive Example 2.5 - Solution (2)

2.	**Name**	**Formula**	**Comments**
a.	Manganese(IV) oxide	MnO_2	Two O^{2-} ions (total charge 4–) are required by the Mn^{4+} ion [manganese(IV)]
b.	Lead(II) chloride	$PbCl_2$	Two Cl^- ions are required by the Pb^{2+} ion [lead(II)] for charge balance

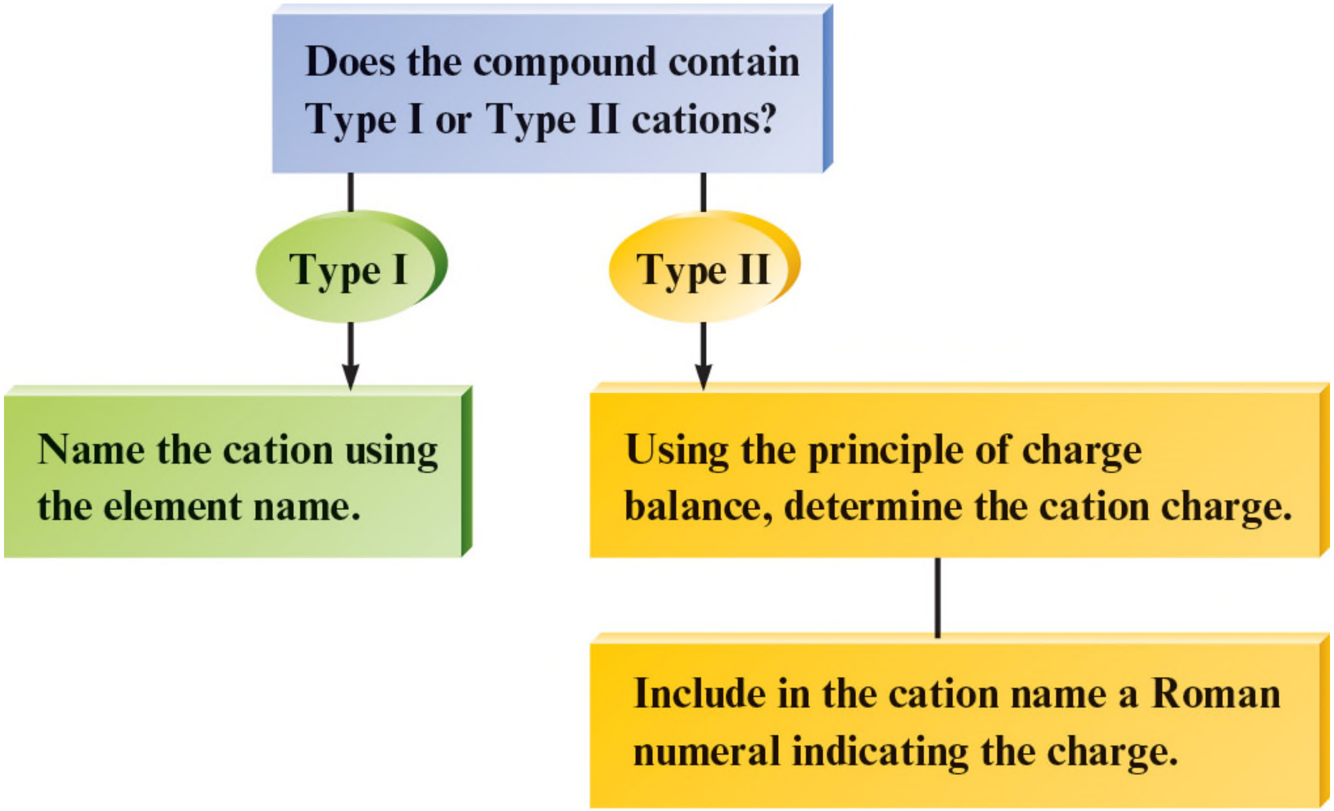
Section 2.8

Naming Simple Compounds

Critical Thinking

- We can use the periodic table to tell us something about the stable ions formed by many atoms
    - For example, the atoms in column 1 always form 1+ ions and the transition metals, however, can form more than one type of stable ion
        - What if each transition metal ion had only one possible charge? How would the naming of compounds be different?

Figure 2.20 - Flowchart for Naming Binary Ionic Compounds



Does the compound contain Type I or Type II cations?

Type I

Name the cation using the element name.

Type II

Using the principle of charge balance, determine the cation charge.

Include in the cation name a Roman numeral indicating the charge.

Section 2.8
Naming Simple Compounds

Interactive Example 2.6 - Naming Binary Compounds

1. Give the systematic name for each of the following compounds:
   a. $CoBr_2$
   b. $CaCl_2$
2. Given the following systematic names, write the formula for each compound:
   a. Chromium(III) chloride
   b. Gallium iodide

Interactive Example 2.6 - Solution (1)

1.	**Formula**	**Name**	**Comments**
a.	$CoBr_2$	Cobalt(II) bromide	Cobalt is a transition metal; the compound name must have a Roman numeral The two Br^- ions must be balanced by a Co^{2+} ion
b.	$CaCl_2$	Calcium chloride	Calcium, an alkaline earth metal, forms only the Ca^{2+} ion A Roman numeral is not necessary

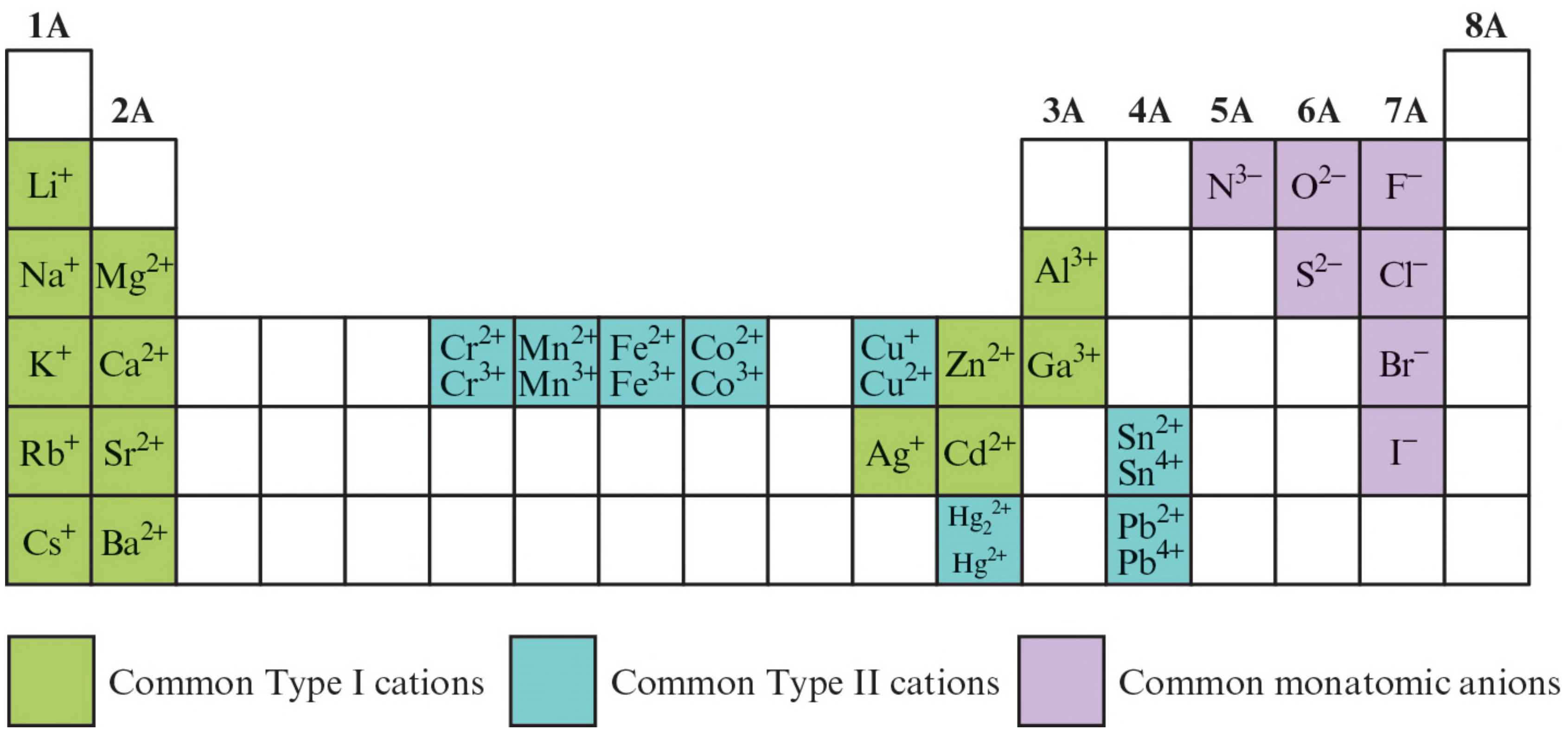
Section 2.8

Naming Simple Compounds

Interactive Example 2.6 - Solution (2)

2.	**Name**	**Formula**	**Comments**
a.	Chromium(III) chloride	$CrCl_3$	Chromium(III) indicates that Cr^{3+} is present, so 3 Cl^- ions are needed for charge balance
b.	Gallium iodide	GaI_3	Gallium always forms 3+ ions, so 3 I^- ions are required for charge balance

Figure 2.21 - Common Cations and Anions



Section 2.8
Naming Simple Compounds

Polyatomic Ions

- Assigned special names that must be memorized for naming compounds
- **Oxyanions**: Anions that contain an atom of a given element and different numbers of O atoms
    - When there are two members in the series:
        - Name of the member with the smaller number of O atoms ends with -ite
        - Name of the member with the larger number of O atoms ends with -ate

Section 2.8

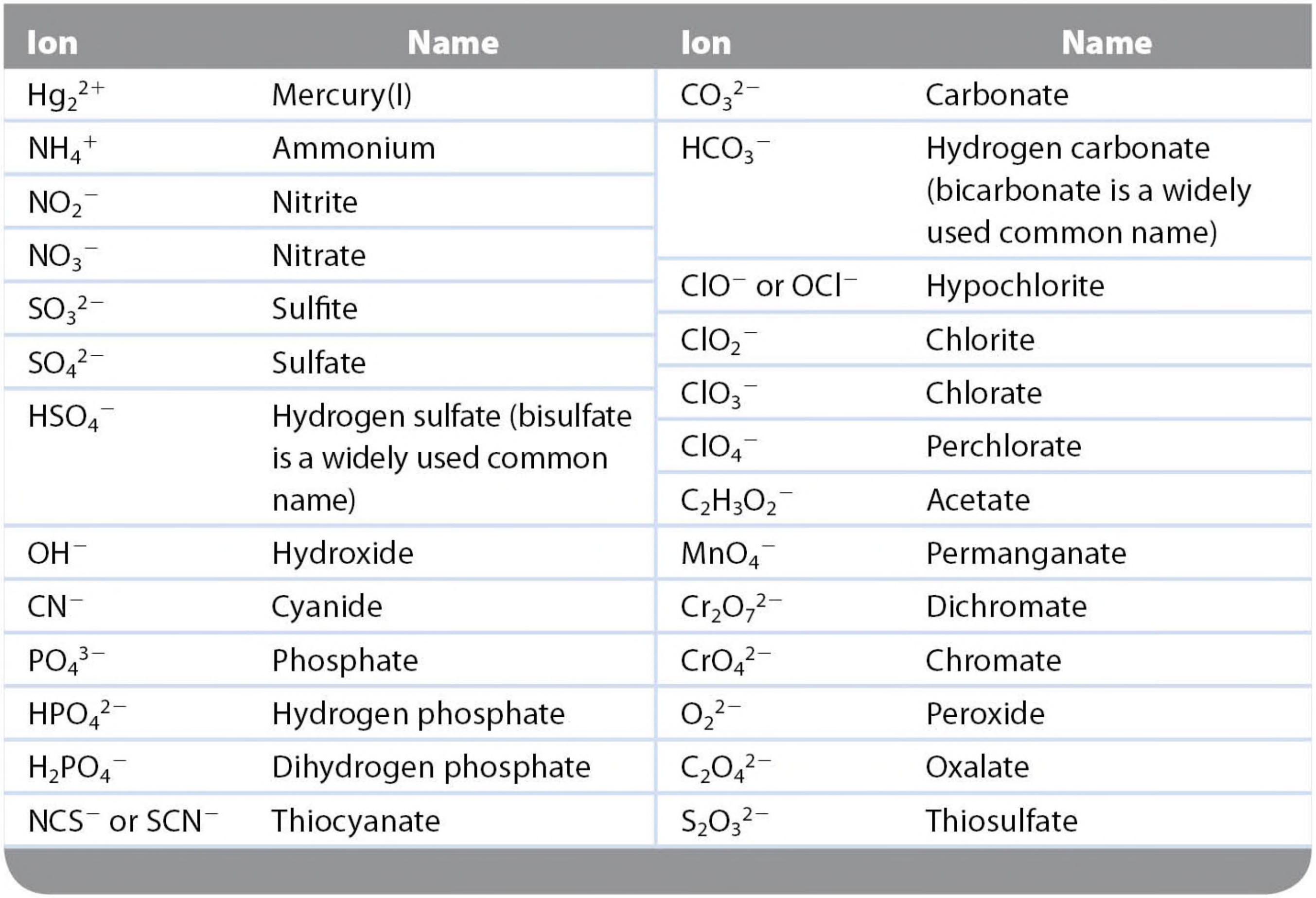
Naming Simple Compounds

Polyatomic Ions (continued)

- When more than two oxyanions make up a series:
  - Use the prefix hypo- (less than) to name members of the series with the fewest O atoms
  - Use the prefix per- (more than) to name members of the series with the most O atoms

Section 2.8
Naming Simple Compounds

Table 2.5 - Common Polyatomic Ions



Ion	Name	Ion	Name
Hg_2^{2+}	Mercury(I)	CO_3^{2-}	Carbonate
NH_4^+	Ammonium	HCO_3^-	Hydrogen carbonate (bicarbonate is a widely used common name)
NO_2^-	Nitrite	ClO^- or OCl^-	Hypochlorite
NO_3^-	Nitrate	ClO_2^-	Chlorite
SO_3^{2-}	Sulfite	ClO_3^-	Chlorate
SO_4^{2-}	Sulfate	ClO_4^-	Perchlorate
HSO_4^-	Hydrogen sulfate (bisulfate is a widely used common name)	$C_2H_3O_2^-$	Acetate
OH^-	Hydroxide	MnO_4^-	Permanganate
CN^-	Cyanide	$Cr_2O_7^{2-}$	Dichromate
PO_4^{3-}	Phosphate	CrO_4^{2-}	Chromate
HPO_4^{2-}	Hydrogen phosphate	O_2^{2-}	Peroxide
$H_2PO_4^-$	Dihydrogen phosphate	$C_2O_4^{2-}$	Oxalate
NCS^- or SCN^-	Thiocyanate	$S_2O_3^{2-}$	Thiosulfate

Section 2.8
Naming Simple Compounds

Interactive Example 2.7 - Naming Compounds Containing Polyatomic Ions

1. Give the systematic name for each of the following compounds:
   a. Na_2SO_4
   b. $Mn(OH)_2$
2. Given the following systematic names, write the formula for each compound:
   a. Sodium hydrogen carbonate
   b. Sodium selenate

Section 2.8
Naming Simple Compounds

Interactive Example 2.7 - Solution (1)

1.	**Formula**	**Name**	**Comments**
a.	Na_2SO_4	Sodium sulfate	
b.	$Mn(OH)_2$	Manganese(II) hydroxide	Transition metal—name must contain a Roman numeral The Mn^{2+} ion balances three OH^- ions

Interactive Example 2.7 - Solution (2)

2.	**Name**	**Formula**	**Comments**
a.	Sodium hydrogen carbonate	$NaHCO_3$	Often called sodium bicarbonate
b.	Sodium selenate	Na_2SeO_4	Atoms in the same group, like sulfur and selenium, often form similar ions that are named similarly Thus SeO_4^{2-} is selenate, like SO_4^{2-} (sulfate)

Binary Covalent Compounds (Type III)

- Formed between two nonmetals
- Naming binary covalent compounds
  - First element in the formula is named first, using the full element name
  - Second element is named as if it were an anion
  - Prefixes are used to denote the numbers of atoms present
  - Prefix mono- is never used for naming the first element

Section 2.8

Naming Simple Compounds

Table 2.6 - Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Interactive Example 2.8 - Naming Type III Binary Compounds

1. Name each of the following compounds:
   a. PCl_5
   b. PCl_3
   c. SO_2
2. From the following systematic names, write the formula for each compound:
   a. Sulfur hexafluoride
   b. Sulfur trioxide

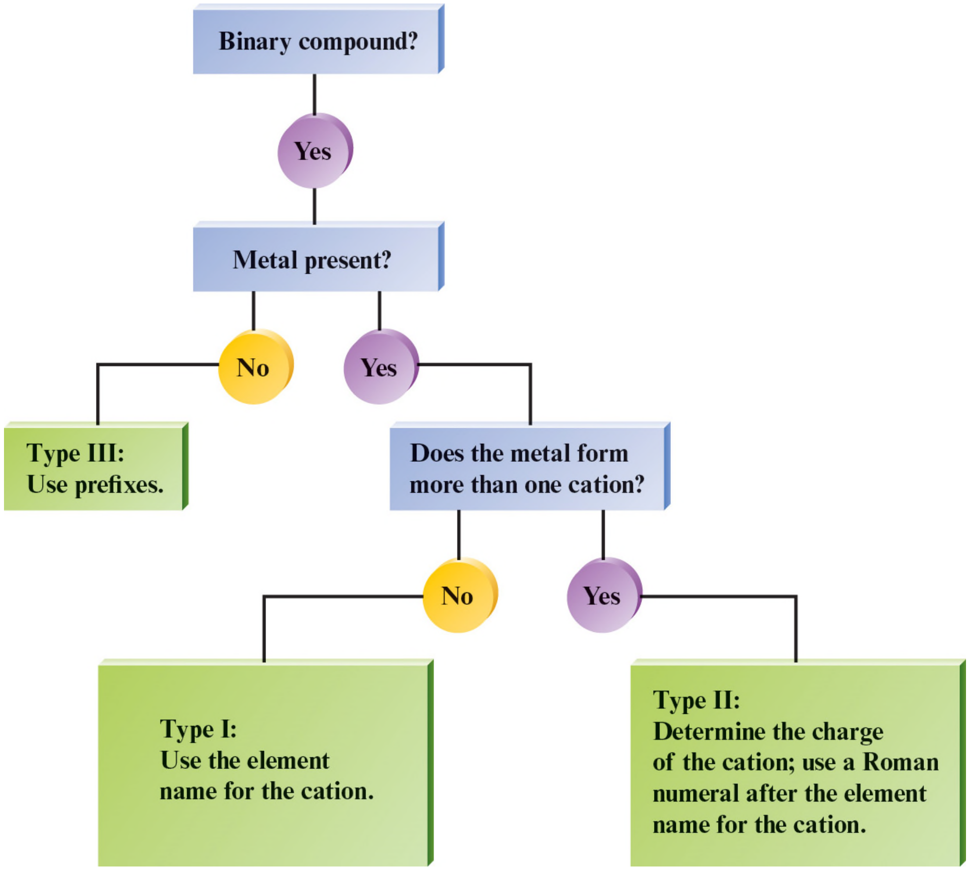
Interactive Example 2.8 - Solution (1)

1.	**Formula**	**Name**
a.	PCl_5	Phosphorus pentachloride
b.	PCl_3	Phosphorus trichloride
c.	SO_2	Sulfur dioxide

Interactive Example 2.8 - Solution (2)

2.	**Name**	**Formula**
a.	Sulfur hexafluoride	SF_6
b.	Sulfur trioxide	SO_3

Figure 2.22 - A Flowchart for Naming Binary Compounds



Binary compound?

Yes

Metal present?

No

Type III:
Use prefixes.

Yes

Does the metal form more than one cation?

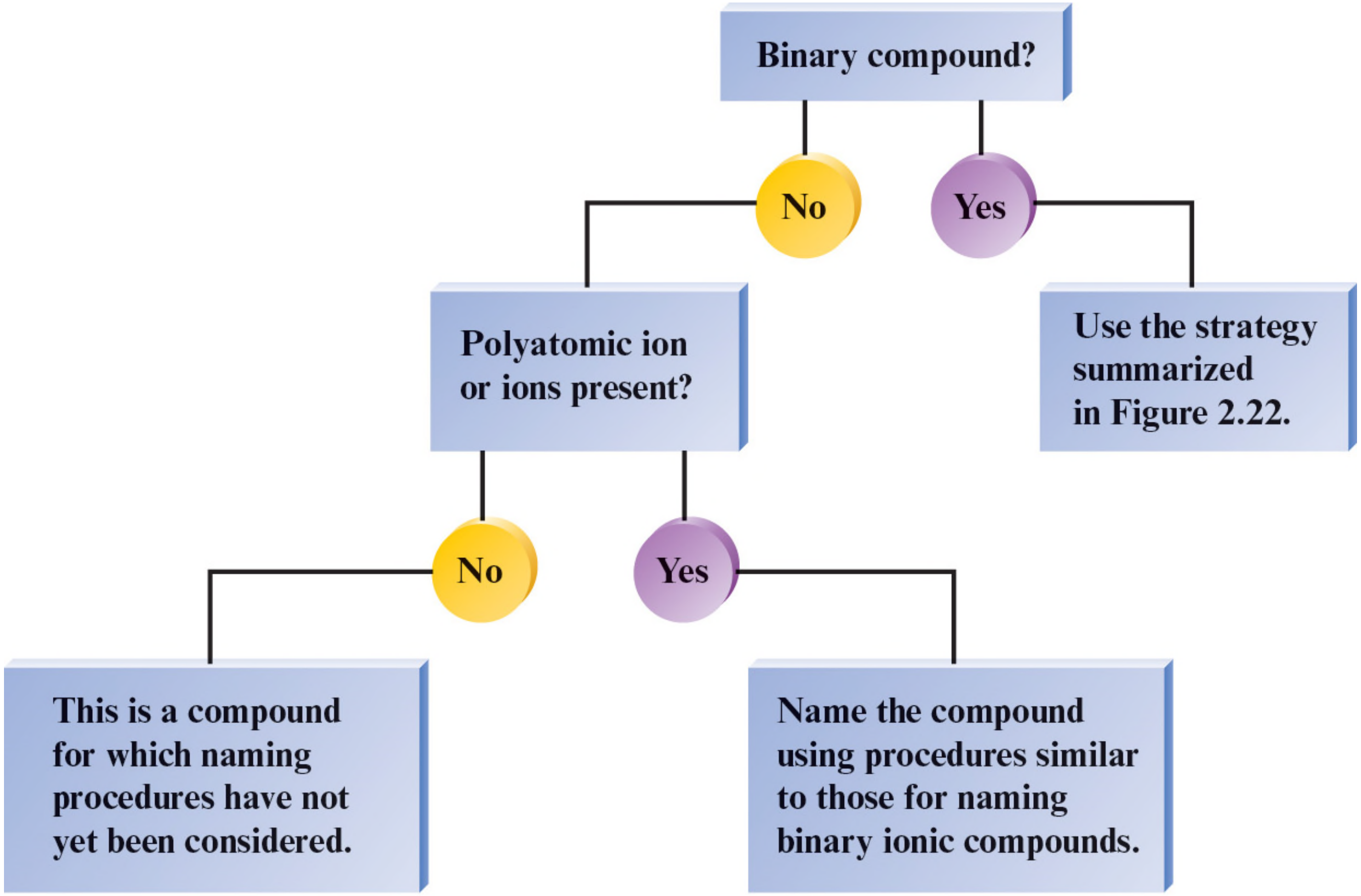
No

Type I:
Use the element name for the cation.

Yes

Type II:
Determine the charge of the cation; use a Roman numeral after the element name for the cation.

Figure 2.23 - Overall Strategy for Naming Chemical Compounds


Binary compound?

No

Yes

Polyatomic ion or ions present?

Use the strategy summarized in Figure 2.22.

No

Yes

This is a compound for which naming procedures have not yet been considered.

Name the compound using procedures similar to those for naming binary ionic compounds.

Acids

- Molecules in which one or more H^+ ions are attached to an anion
- Nomenclature rules depend on whether the anion contains oxygen
  - If the anion ends in -ide, the acid is named with the prefix hydro- and the suffix -ic
  - When the anion contains oxygen, based on the name of the anion, the acidic name is formed from the root name of the anion with a suffix of -ic or -ous

Section 2.8

Naming Simple Compounds

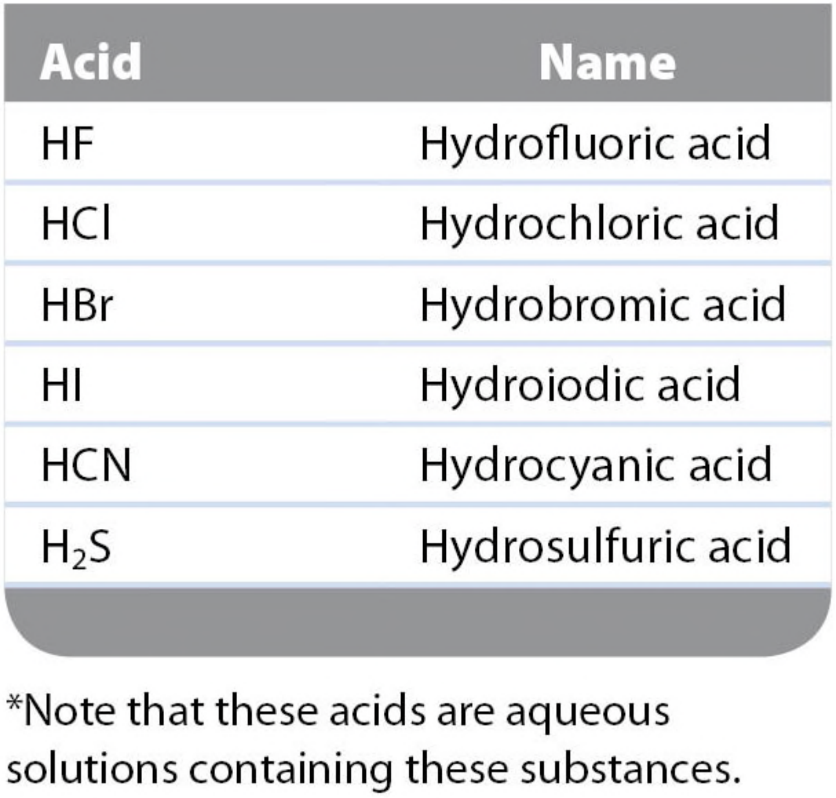
Naming Acids When the Anion Contains Oxygen

- If the anion name ends in -ate, the suffix -ic is added to the root name
- If the anion has an -ite ending, the -ite is replaced by -ous

Section 2.8

Naming Simple Compounds

Table 2.7 - Names of Acids* That Do Not Contain Oxygen

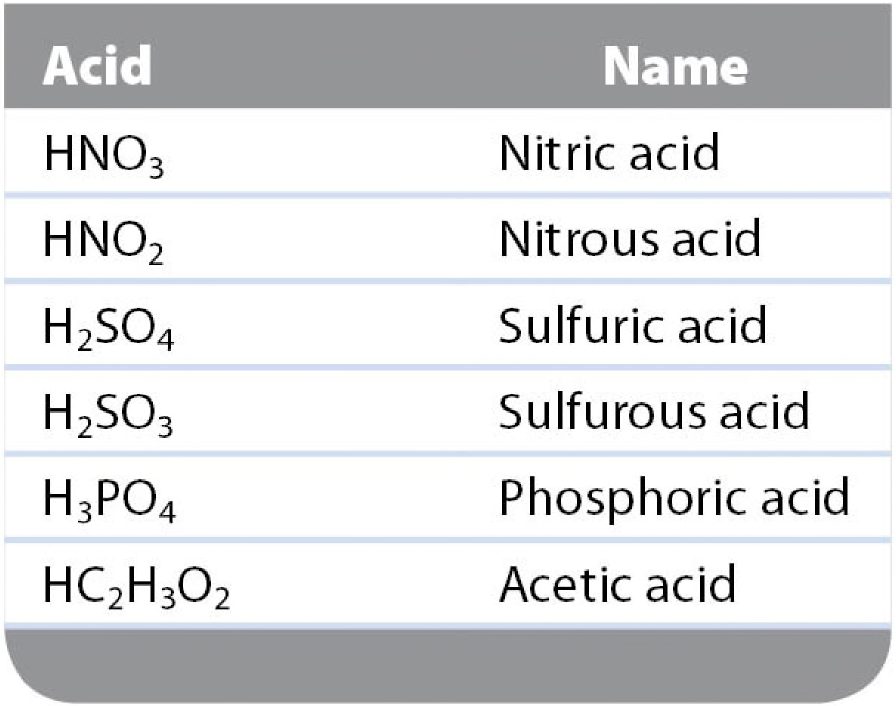


Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
HCN	Hydrocyanic acid
H_2S	Hydrosulfuric acid

*Note that these acids are aqueous solutions containing these substances.

Section 2.8
Naming Simple Compounds

Table 2.8 - Names of Some Oxygen-Containing Acids



Acid	Name
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic acid

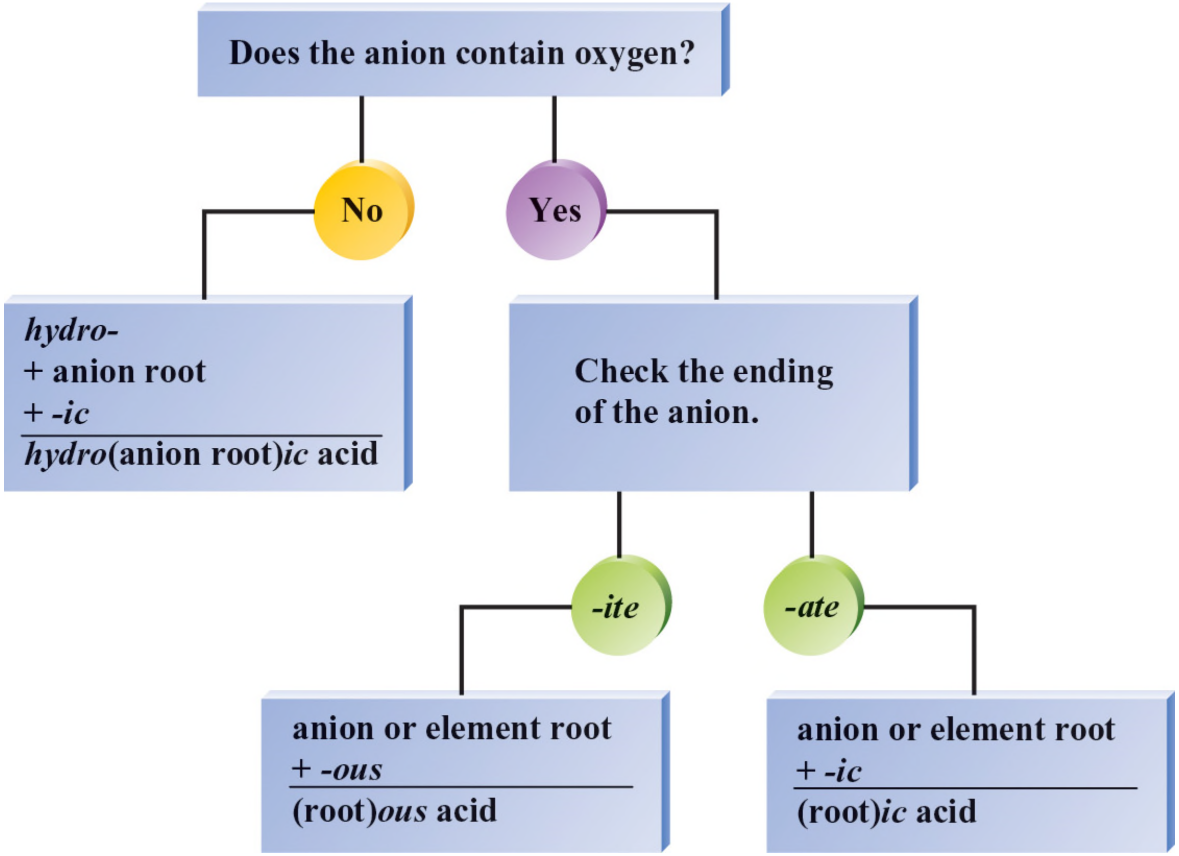
Critical Thinking

- In this chapter, you have learned a systematic way to name chemical compounds
- What if all compounds had only common names?
    - What problems would this cause?

Section 2.8

Naming Simple Compounds

Figure 2.24 - A Flowchart for Naming Acids



Does the anion contain oxygen?

No

hydro-
+ anion root
+ *-ic*

hydro(anion root)*ic* acid

Yes

Check the ending of the anion.

-ite

anion or element root
+ *-ous*

(root)*ous* acid

-ate

anion or element root
+ *-ic*

(root)*ic* acid

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Exercise

- Name each of the following compounds:

a.	CuI	**Copper(I) iodide**
b.	S_4N_4	**Tetrasulfur tetranitride**
c.	$NaHCO_3$	**Sodium hydrogen carbonate or sodium bicarbonate**
d.	$BaCrO_4$	**Barium chromate**

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Join In (8)

- Which is not the correct chemical formula for the compound named?
  a. Potassium phosphate, K_3PO_4
  b. Iron(II) oxide, FeO
  c. Calcium carbonate, $CaCO_3$
  d. Sodium sulfide, NaS
  e. Lithium nitrate, $LiNO_3$

Join In (9)

- Which of the following is named incorrectly?
  a. Li_2O, lithium oxide
  b. $FePO_4$, iron(III) phosphate
  c. HF, hydrogen fluoride
  d. $BaCl_2$, barium dichloride
  e. Mg_3N_2, magnesium nitride

Join In (10)

- Which is the correct formula for copper(I) sulfide?
  a. CuS
  b. Cu_2S
  c. CuS_2
  d. Cu_2S_2
  e. None of these

Join In (11)

- Which of the following is the correct chemical formula for iron(III) oxide?
  a. FeO
  b. Fe_3O
  c. FeO_3
  d. Fe_2O_3
  e. Fe_3O_2

Join In (12)

- What is the correct name for the compound with the formula $Mg_3(PO_4)_2$?
  a. Trimagnesium diphosphate
  b. Magnesium(II) phosphate
  c. Magnesium phosphate
  d. Magnesium(II) diphosphate
  e. Magnesium(III) diphosphate

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Join In (13)

- What is the correct name for the acid with the formula HFO?
  a. Fluoric acid
  b. Hydrofluoric acid
  c. Hydrofluorous acid
  d. Hypofluorous acid
  e. Perfluoric acid