Chapter 2

Atoms, Molecules, and Ions

OVERVIEW OF THE CHAPTER

2.1, 2.2, 2.3 ATOMS

Learning Goals: You should be able to:

1. Describe the composition of an atom in terms of protons, neutrons, and electrons.
2. Give the approximate size, relative mass, and charge of an atom, proton, neutron, and electron.
3. Write the chemical symbol for an element, having been given its mass number and atomic number, and perform the reverse operation.
4. Describe the properties of the electron as seen in cathode rays. Describe the means by which J.J. Thomson determined the ratio *e/m* for the electron.
5. Describe Millikan's oil-drop experiment and indicate what property of the electron he was able to measure.
6. Cite the evidence from studies of radioactivity for the existence of subatomic particles.
7. Describe the experimental evidence for the nuclear nature of the atom.

2.4, 2.5 PERIODIC TABLE: ARRANGEMENT OF ATOMS AND ATOMIC WEIGHTS

Learning Goals: You should be able to:

1. Use the unit of atomic mass unit (amu) in calculation of masses of atoms.
2. Define the term atomic weight and calculate the atomic weight of an element given its natural distribution of isotopes and isotopic masses.
3. Use the periodic table to determine the atomic number, atomic symbol, and atomic weight of an element.
4. Define the terms group and period and recognize the common groups of elements.
5. Use the periodic table to predict whether an element is metallic, nonmetallic, or metalloid.

2.6, 2.7 MOLECULES, MOLECULAR COMPOUNDS, IONIC COMPOUNDS, AND IONS

Learning Goals: You should be able to:

1. Define the term molecule and recognize which elements typically combine to form molecules.
2. Distinguish between empirical and molecular formulas.
3. Draw the structural and ball-and-stick formulas of a substance given its chemical formula and the linkage between atoms.

4. Use the periodic table to predict the charges of monatomic ions of non-transition elements.
5. Write the symbol and charge for an atom or ion having been given the number of protons, neutrons, and electrons and perform the reverse operation.
6. Determine whether a substance is likely to be ionic or molecular.
7. Write the simplest formula of an ionic compound having been given the charges of ions from which it is made.

2.8, 2.9 NOMENCLATURE: SIMPLE INORGANIC, MOLECULAR, AND ORGANIC COMPOUNDS

Learning Goals: You should be able to:

1. Write the name of a simple inorganic compound having been given its chemical formula and perform the reverse reaction.
2. Write and name common polyatomic ions.
3. Write and name acids based on anions whose names end in -ide, -ate, and -ite.
4. Write the name of simple binary molecular compounds and perform the reverse operation.
5. Define the terms hydrocarbon, alkane, and alcohol and be able to name simple alkanes and alcohols, having been given the chemical formula, and perform the reverse operation.

TOPIC SUMMARIES AND EXERCISES

ATOMS

This chapter focuses on the structure of atoms, elements in the periodic table, how atoms combine to form substances, and naming inorganic compounds. The work of many scientists, such as John Dalton, J.J. Thompson, and R. Millikan, provided experimental information for development of a modern understanding of the structure of atoms. You should know the contributions of the key historical figures discussed in this chapter: See Exercise 2.

Atoms are the smallest particles of an element that have all the same chemical properties of that element. Some general characteristics about the structure of atoms are:

- The mass of a single atom is extremely small, less than 10^{-21} g.
- The center of an atom contains a small **nucleus**. A nucleus contains **protons** and **neutrons**, which make up most of the mass of an atom.
- The volume of an atom mostly consists of **electrons**.
- Protons (positively charged) and neutrons (no charge) have essentially the same mass. Electrons (negatively charged) are significantly less massive.

All atoms of the same element have the same number of protons.

- Atoms that have the same number of protons but differ in their number of neutrons are called **isotopes**.
- The general symbol for an isotope is

$^{A}_{Z}X$
- **Mass number** = number of protons + number of neutrons
- Symbol of element
- **Atomic number** = number of protons

Note: The atomic number Z defines an element; it also tells you how many electrons a neutral atom of that element possesses. Thus, the number of neutrons equals A - Z.

EXERCISE 1 Writing isotopic symbols

Write the nuclear-isotope symbols for the four isotopes of sulfur with 16, 17, 18, and 20 neutrons, respectively.

SOLUTION: The atomic number of sulfur is 16, and all of its isotopes will have 16 protons. The mass number of each isotope is the sum of its number of neutrons plus its number of protons: $16 + 16 = 32$; $16 + 17 = 33$; $16 + 18 = 34$; and $16 + 20 = 36$. The nuclear-isotope symbols are $^{32}_{16}S$, $^{33}_{16}S$, $^{34}_{16}S$, and $^{36}_{16}S$. Note that the subscript before each elemental symbol (that is, the atomic number) is the same for all isotopes because the number of protons is invariant for sulfur.

EXERCISE 2 Developments in the history of atomic structure

The composition of the atom was determined from the experiments of many individuals. For each experiment, state the principal person associated with the experiment, the particle studied, and the property of the particle determined: **(a)** α scattering; **(b)** cathode ray; **(c)** oil drop.

SOLUTION: **(a)** E. Rutherford determined the relative volume and mass of the nucleus in the atom by observing the scattering of α particles. **(b)** J. J. Thomson determined the *e/m* ratio of the electron by using a cathode-ray tube. **(c)** R. Millikan determined the charge of an electron by observing the motion of charged oil drops in an electric field.

EXERCISE 3 Using isotopic symbols to determine nuclear composition

Find the number of protons, electrons, and neutrons in the following isotopes:

(a) $^{40}_{20}Ca$; **(b)** $^{238}_{92}U$.

SOLUTION: **(a)** The number of protons is the subscript number shown in the isotopic symbol: 20. The number of electrons in a neutral atom is the same as the number of protons: 20. The number of neutrons equals the mass number minus the number of protons. This is the superscript number (mass number) minus the subscript number: $40 - 20 = 20$. **(b)** Preceding as in **(a)**, the number of protons is 92, the number of electrons is 92, and the number of neutrons is $238 - 92 = 146$.

EXERCISE 4 Changing the number of electrons in an isotope

If an electron is added or removed from ^{35}Cl, what changes occur to the isotope?

SOLUTION: The isotope is still ^{35}Cl because the number of protons and the number of neutrons do not change; however, the number of electrons does. If electrons, which are negatively charged, are added to an isotope, the isotope gains negative charge. Therefore if one electron is added to ^{35}Cl, it becomes $^{35}Cl^-$. Similarly, if an electron is removed, the isotope loses negative charge and becomes positively charged because there are more protons than electrons: $^{35}Cl^+$.

PERIODIC TABLE: ARRANGEMENT OF ATOMS AND ATOMIC WEIGHTS

The masses of isotopes are measured relative to the atomic mass of $^{12}_{6}C$, which is defined to have a mass of exactly 12 **atomic mass units** (amu). A modern mass spectrometer is used to determine accurate relative masses.

- Except for a few elements, most elements have a distribution of naturally occurring isotopes. The concept of atomic weight is based on determining the average of the masses of naturally occurring isotopes weighted according to their abundances:

$$\text{atomic weight} = \%\ \text{abundance} \times \text{mass of isotope 1} + \%\ \text{abundance} \times \text{mass of isotope 2} + \cdots + \%\ \text{abundance} \times \text{mass of isotope n.}$$

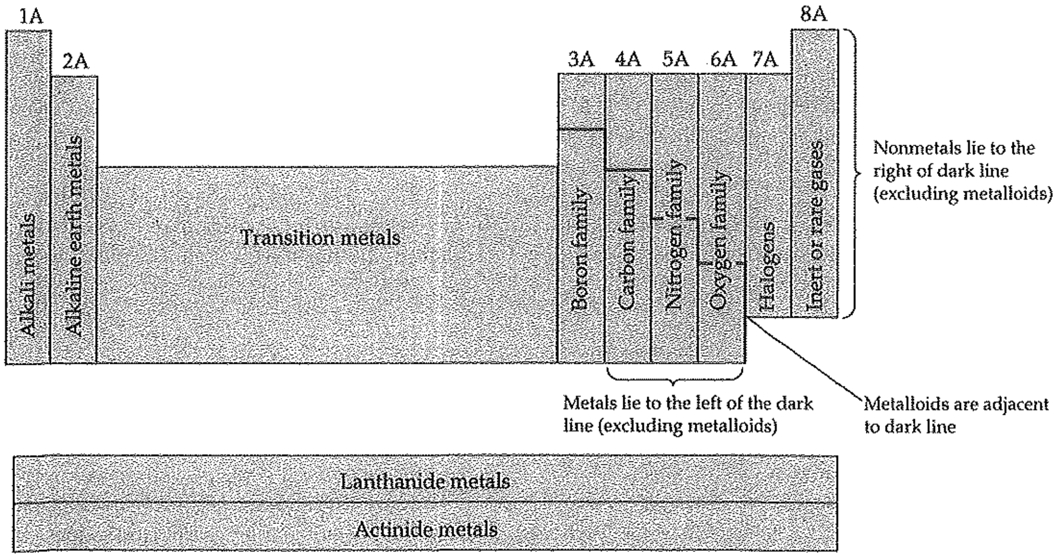
- The relationship between amu and grams is 1.66054×10^{-24} g = 1 amu.

The periodic table lists elements in order of increasing atomic number. It is very useful in helping us to organize trends in the chemical and physical properties of elements. The manner in which elements are grouped in vertical columns (families) and horizontal rows (periods) helps you to remember that

- *Metals* are on the left side and middle (except hydrogen).
- *Nonmetals* are on the right side.
- *Metalloids* have properties of both metals and nonmetals. They are identified by the dark diagonal line toward the right side of the periodic table.
- Elements in vertical columns, called *families* or *groups*, exhibit similar chemical and physical properties, whereas elements in a horizontal row (*period*) exhibit different properties. The names of groups in the periodic table and the general location of metals, metalloids, and nonmetals are shown in Figure 2.1.

Exercise 5 Understanding the concept of average atomic weight

(a) What is the sum of percentages of naturally occurring isotopes for an element?
(b) The following problem shows the concept of atomic weights applied to a different situation. What is the average weight of a class of students given the following information about the class: (1) The average woman weighs 122 pounds; (2) the average man weighs 165 pounds; and (3) the percentage of men in the class is 45.0%?



▲ **FIGURE 2.1** Families in the periodic table. Note that hydrogen, the first element in family 1A, is not an alkali metal.

SOLUTION: (a) The sum of percentages must add to 100%; however, in experimental situations it rarely adds exactly to this value because of experimental error. (b) The average atomic weight of the class is calculated from the relation:

$$\begin{aligned}\text{average weight} &= \%\text{ men} \times \text{average mass of men} \\ &\quad + \%\text{ women} \times \text{averagemass of women}\end{aligned}$$

We are given the percentage of men but not that of women. We can calculate the percentage of women from relation:

$$100.0\% = \%\text{ men} + \%\text{ women} = 45.0\% + \%\text{ women}$$

$$\%\text{ women} = 100\% - 45.0\% = 55.0\%$$

Therefore, the average weight of the class is:

$$\text{average weight} = (0.450)(165\text{ pounds}) + (0.550)(122\text{ pounds}) = 141\text{ pounds}$$

EXERCISE 6 Calculating the atomic weight of an element

What is the atomic weight of antimony if it has only two naturally occurring isotopes, Sb-121 with an isotopic mass of 120.904 amu and an abundance of 57.21% and Sb-123 with an isotopic mass of 122.904 amu and an abundance of 42.79%?

SOLUTION: We can use the concept of average atomic mass discussed in the previous exercise to solve for the atomic weight of antimony:

$$\begin{aligned}\text{Atomic weight of Sb} &= 120.904\text{ amu}(0.5721) + 122.904\text{ amu}(0.4279) \\ &= 69.17\text{ amu} + 52.59\text{ amu} \\ &= 121.76\text{ amu}\end{aligned}$$

EXERCISE 7 Using a periodic table

The position of an element in the periodic table is a function of its atomic composition. Elements are arranged in order of increasing atomic number. The atomic number of an atom equals its number of protons. By referring to a periodic table, determine which *neutral* atom: (a) contains 50 protons; (b) contains 17 electrons; (c) has an atomic number of 56.

SOLUTION: (a) Tin (Sn). Its atomic number, 50, corresponds to 50 protons. (b) Chlorine (Cl). Its atomic number corresponds to 17 protons. This is also the number of electrons in the neutral atom. (c) Barium (Ba). Its atomic number is 56.

EXERCISE 8 Identifying the family of an element in the periodic table

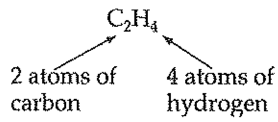
What is the name of the family in which the following elements occur: I, Na, S, and Ca?

SOLUTION: By referring to Figure 2.1 and a periodic table we find the requested information: I, halogen; Na, alkali metal; S, oxygen family; and Ca, alkaline-earth metal.

MOLECULES, MOLECULAR COMPOUNDS, IONIC COMPOUNDS, AND IONS

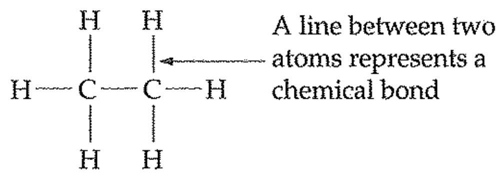
Molecular compounds consist of small particles called molecules.

- A **molecule** is a small particle consisting of two or more atoms combined together in a discrete unit. Molecules have their own chemical and physical properties, which differ from the properties of the elements forming them.
- Molecules typically consist of nonmetallic elements.
- You should know the seven elements that form homonuclear diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.
- Subscripts in a **molecular formula** tell you how many atoms are actually present



- An **empirical formula** shows only the simplest whole number ratio of atoms. For example, C_2H_4 is a molecular formula; its empirical formula, CH_2, is obtained by dividing the subscripts in the molecular formula by 2: $CH_2 = C_{2/2}H_{4/2}$
- A **structural formula** shows "how atoms are joined together"—that is, the relative orientation and position of bonded atoms. For example, the structural formula of C_2H_6 shows the following arrangement of atoms:

An **ionic compound** consists of positively charged ions (cations) and negatively charged ions (anions).



- Ionic compounds typically contain a metal and a nonmetal.
- **Cations** are positively charged ions of metallic elements formed by the loss of electrons; e.g., Na^+ is formed by Na losing one electron.
- **Anions** are negatively charged ions of nonmetallic elements formed by the addition of electrons; e.g., S^{2-} is formed by S gaining two electrons.
- Ionic compounds do not contain unique identifiable pairs of ions. Therefore, empirical formulas are used to describe their composition.
- Ionic compounds may contain polyatomic ions such as NO_3^- as in $NaNO_3$. Polyatomic ions are like molecules except that they carry a charge.
- You should know the most common charge carried by elements in the following families:

Alkali metals	Alkaline-earth metals	Transition metals	Halogens	Oxygen family
1+	2+	2+, 3+	1−	2−

EXERCISE 9 Interpreting chemical formulas

Interpret the following chemical formulas: (a) N_2; (b) NH_3; (c) NH_4^+; (d) H_2SO_4.

SOLUTION: (a) N_2 is a homonuclear (having atoms of the same type) molecule composed of two nitrogen atoms; this is the elemental form of nitrogen. (b) NH_3 is a heteronuclear (having more than one type of atom) molecule composed of one nitrogen atom and three hydrogen atoms. (c) NH_4^+ is an ion that has a 1+ charge and is composed of one nitrogen atom and four hydrogen atoms. (d) H_2SO_4 is a heteronuclear molecule composed of two hydrogen atoms, one sulfur atom, and four oxygen atoms.

EXERCISE 10 Writing empirical and chemical formulas

(a) Do C_2H_4 and 2 CH_2 represent the same substance? (b) What are the empirical formulas for B_2H_6, H_2O, and $C_2H_2O_4$?

SOLUTION: (a) The formulas do not represent the same substance. C_2H_4 represents the molecular formula for the substance acetylene, a gas used in welding. 2 CH_2 represents two empirical formulas of acetylene. (b) To determine an empirical formula, divide each subscript of a molecular formula by the largest whole number that goes into each subscript (the greatest common denominator)

$$B_{2/2}H_{6/2} = BH_3$$

There is no way to further reduce the subscripts in H_2O—this is the molecular and empirical formula for water.

$$C_{2/2}H_{2/2}O_{4/2} = CHO_2$$

EXERCISE 11 Writing chemical formulas of ionic substances

Write the formula for: (a) a neutral polyatomic compound consisting of one barium ion with a 2+ charge (written Ba^{2+}) and chlorine ions with a 1− charge (written Cl^-). (b) A polyatomic ion that has a 1− charge and consists of one boron ion with a 3^+ charge (written B^{3+}) and fluorine ions with a 1− charge (written F^-).

SOLUTION: (a) The sum of positive and negative charges for the ions must equal zero in a neutral compound. Two Cl^- ions are required to balance the 2+ charge of Ba^{2+}: $BaCl_2$. (b) The sum of positive and negative charges for the ions must equal the 1− charge of the compound. Four F^- ions are required so that, when they are combined with B^{3+}, the sum of the charges is 1−: $BF_4^-[+3 + 4(-1) = -1]$.

EXERCISE 12 Determining the number of electrons possessed by a molecule or ion

Determine the total number of electrons in each of the following: (a) NH_3; (b) NH_4^+; (c) NH_2^-.

SOLUTION: Determine the number of electrons for each atom in each chemical formula, sum these numbers, and then adjust the number for the charge shown. For each negative charge, one electron is added; for each positive charge, one electron is subtracted. (a) The number of electrons for each atom is the same as the number of protons. Each hydrogen atom has one proton and each nitrogen atom has 7 protons. Thus, for NH_3 the number of electrons is $(1 \times 7) + (3 \times 1) = 10$ electrons. The charge of the compound is zero, thus no adjustment for charge is necessary. (b) The number of electrons for NH_4^+ with no charge is $(1 \times 7) + (4 \times 1) = 11$ electrons.

However, the species has a positive charge, thus one electron must be removed so that there are more protons than electrons: $11 - 1 = 10$ electrons for NH_4^+. (c) The number of electrons for NH_2^- with no charge is $(1 \times 7) = (2 \times 1) = 9$ electrons. The species has a charge of 1−, thus one electron must be added so that there are more electrons than protons: $9 + 1 = 10$ electrons.

EXERCISE 13 Writing binary chemical formulas

Write the expected formula when the following elements combine to form compounds: (a) Al and O; (b) B and Cl; (c) Ca and F; (d) Na and I; (e) S and O; and (f) H and Ca.

SOLUTION: (a) Al forms the 3+ state in salts. Oxygen forms the oxide ion, O^{2-}: Al_2O_3. (b) Boron is a member of the family containing aluminum; thus, you should expect it to behave in a similar manner to aluminum. BCl_3. (c) Calcium forms Ca^{2+} and flourine forms F^-: CaF_2. (d) Na forms Na^+ and iodine in the presence of metals forms I^-: NaI. (e) Sulfur and oxygen, both nonmetals, form nonmetal oxides: For example, $SO_2(g)$ and $SO_3(g)$. (f) Ca is an active metal. In the presence of an active metal hydrogen forms the hydride ion, H^-: CaH_2.

NOMENCLATURE: SIMPLE INORGANIC, MOLECULAR, AND ORGANIC COMPOUNDS

Chemists not only write chemical formulas for compounds but also give them reasonably systematic names. Rules for naming simple inorganic and binary molecular compounds are found in Section 2.8 in the text. *These rules should be committed to memory; and you should then practice them by naming anions, cations, ionic compounds, oxyanions, acids, and simple binary molecular compounds.* Attempt Exercises 13–19 *after* you have reviewed nomenclature rules. The solutions to many of the exercises provide explanations which reinforce what you have learned.

In Section 2.9 the topic of organic chemistry is introduced with a focus on hydrocarbons and the alcohol functional group. *Hydrocarbons* are compounds containing hydrogen and carbon atoms.

- One class of hydrocarbons is called an *alkane*, in which each carbon atom is bonded to four other atoms; all bonds are single. You need to learn the prefixes for alkanes in Table 2.6 in the text.
- A *functional group* is a nonhydrogen atom or group of atoms that replaces a hydrogen atom in a hydrocarbon. *Alcohols* contain the – OH functional group and their names end in *-ol*.

EXERCISE 14 Correcting chemical formulas when they are written incorrectly

The following chemical formulas are written incorrectly. Correct them so that they are in the proper form: (a) ClNa; (b) CaClCl; (c) $NH_4NH_4SO_4$; (d) H^2S; (e) FXeF.

SOLUTION: (a) NaCl. The cation (Na^+) is always placed before the anion (Cl^-). (b) $CaCl_2$. A subscript is used to indicate the number of Cl^- ions present. (c) $(NH_4)_2SO_4$. If there is more than one type of polyatomic ion present, its formula is enclosed in parentheses, and subscripts are used as necessary. (d) H_2S. The 2 must be a subscript when it indicates the number of the same type of atoms or ions present. (e) XeF_2.

EXERCISE 15 Naming cations

Name the following cations: (a) H^+; (b) Na^+; (c) Be^{2+}; (d) Al^{3+}; (e) Co^{2+}; (f) Co^{3+}.

SOLUTION: Monatomic cations have the same names as the elements from which they are formed. Some monatomic cations commonly exist in two common ion states; for example, Fe^{2+} and Fe^{3+}. When naming such ions, indicate the charge of the ion in Roman numerals enclosed within parentheses after the name of the element. Sometimes an older method is used, in which the suffix *-ic* indicates the higher charge and *-ous* the lower one. This method is further complicated by the fact that the Latin name of the element is often used. Thus Fe^{2+} is iron(II) or ferrous, and Fe^{3+} is iron(III) or ferric. When in doubt, use the newer method. (a) Hydrogen. (b) Sodium. (c) Beryllium. (d) Aluminum. (e) Cobalt(II) or cobaltous. (f) Cobalt(III) or cobaltic.

EXERCISE 16 Naming anions

Name the following anions: (a) F^-; (b) S^{2-}; (c) OH^-; (d) CN^-; (e) PO_4^{3-}; (f) PO_3^{3-}; (g) IO^-; (h) IO_2^-; (i) IO_3^-; (j) IO_4^-; (k) HS^-; (l) $H_2PO_4^-$.

SOLUTION: (a) Fluoride. Monatomic anions have *-ide* added to the stem of the element's name. (b) Sulfide. (c) Hydroxide. Treated as a monatomic anion. (d) Cyanide. Treated as a monatomic anion. (e) Phosphate. When two oxyanions exist, as in the cases of PO_4^{3-} and PO_3^{3-}, the one with more oxygen atoms ends in *-ate*. (f) Phosphite. The one with fewer oxygen atoms ends in *-ite*. (g) Hypoiodite. The prefix *hypo-* indicates one less oxygen atom than in IO_2^-. (h) Iodite. (i) Iodate. (j) Periodate. The prefix *per-* indicates one more oxygen atom than in IO_3^-. (k) Hydrogen sulfide. *Hydrogen* is added to the name of the anion to indicate that there is one hydrogen atom present. (l) Dihydrogen phosphate. *Dihydrogen* is added to phosphate to indicate that there are two hydrogen atoms present.

EXERCISE 17 Naming ionic and molecular compounds

Name the following compounds: (a) ZnS; (b) $(NH_4)_2SO_4$; (c) $FeCl_2$; (d) $KClO_4$; (e) $SnCl_2$; (f) PCl_3; (g) SF_6; (h) CO.

SOLUTION: Remember: When naming ionic compounds cations are named before anions; in molecular compounds the atoms are named in the order found in the molecular formula. For binary molecular compounds, the first atom is given its elemental name, and the second is named as if it were a negative ion. (a) Zinc sulfide. Zinc exists only in the 2+ state. (b) Ammonium sulfate. The use of *di-* before *ammonium* to indicate two NH_4^+ ions is not necessary because the −2 charge of SO_4^{2-} automatically requires two NH_4^+ ions. (c) Iron(II) chloride or, using the older method, ferrous chloride. (d) Potassium perchlorate. (e) Tin(II) chloride or stannous chloride. (f) Phosphorus trichloride. A compound formed from two nonmetallic elements is named as if it were an ionic compound, with the appropriate prefixes added to indicate how many atoms are present. (g) Sulfur hexafluoride. The prefix *hexa-* indicates six fluorine atoms. (h) Carbon monoxide.

EXERCISE 18 Naming acids

Name the following acids in water: (a) HCN; (b) H_2S; (c) H_2CO_3; (d) H_3PO_4.

SOLUTION: (a) Hydrocyanic acid. For hydrogen acids formed from monatomic anions, *hydro-* is added to the name of the anion—in this case, *cyanide*—and the *-ide* ending of the anion is changed to an *-ic* ending. (b) Hydrosulfuric acid. Again, the prefix *hydro-* indicates a hydrogen acid, and *-ic* is added to the end of sulfur. (c) Carbonic acid. When an acid is derived from a polyatomic anion whose name ends in *-ate*—*carbonate* (CO_3^{2-}) in this case—the ending *-ic* is added to the name of the central atom of the polyatomic anion, and no prefix is used. (d) Phosphoric acid. Its name is derived from the polyatomic anion *phosphate* (PO_4^{3-}). The *-ate* ending is dropped and is changed to *-oric*. The *-ic* ending for acids is associated with acids formed from anions ending in *-ate*.

EXERCISE 19 Naming molecular compounds

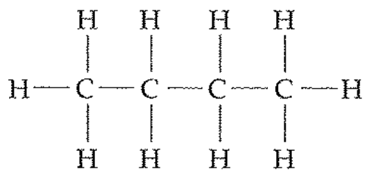
Name the following substances: (a) SF_6; (b) NO; (c) N_2O; (d) P_2O_5.

SOLUTION: We will use the prefixes in Table 2.6 of the text to name the substances. (a) sulfur hexaflouride; (b) nitrogen monoxide; (c) dinitrogen monoxide; (d) diphosphorus pentaoxide.

EXERCISE 20 Writing names and chemical formulas of alkanes

Consider the alkane called butane. (a) Assuming the carbon atoms are in a straight chain, write a structural formula for butane. (b) What is its molecular formula? (c) If one of the hydrogen atoms attached to a carbon atom at the end of the chain is replaced with – OH, what is its name?

SOLUTION: (a) The ending *-ane* in propane tells you that it is an alkane. Referring to Table 2.6 in the text we read that the prefix *but-* means four carbon atoms. We can write



(b) With the structural formula written we can determine the molecular formula by counting the number of carbon and hydrogen atoms: C_4H_{10}. (c) If one of the hydrogen atoms at the end of the carbon chain is replaced by – OH, an alcohol functional group, the following alcohol is produced: Butanol.

SELF-TEST QUESTIONS

Key Terms

Having reviewed key terms in Chapter 2, match key terms with phrases and identify statements as true or false. If a statement is false, indicate why it is incorrect.

Match each phrase with the best term:

2.1 A general name for a positive ion.
2.2 A sulfide ion is an example of this species.
2.3 The least massive particle in an atom as discussed in the chapter.
2.4 A particle possessing no charge in the nucleus.
2.5 Helium atom with a +2 charge.
2.6 $^{85}_{37}Rb$ and $^{87}_{37}Rb$ are examples.
2.7 The number 85 in $^{85}_{37}Rb$.
2.8 The number 37 in $^{85}_{37}Rb$.
2.9 HCl is an example of this species.
2.10 Occupies a small volume of an atom.
2.11 NO_3^- is an anion and also called this type of particle.
2.12 Shows smallest whole-number ratio of atoms.
2.13 Shows arrangement of atoms and bonds.
2.14 Name of any type of charged atom.
2.15 H_2O_2 is an example of this type of chemical formula.
2.16 Contains a beam of electrons.
2.17 Contains two atoms bonded together which act as a unit.
2.18 Emits gamma rays.
2.19 Particles of an element having the same atomic number.

Terms:

(a) alpha particle
(b) anion
(c) atomic number
(d) atoms
(e) cation
(f) cathode ray
(g) diatomic molecule
(h) electron
(i) empirical formula
(j) ion
(k) isotopes
(l) mass number
(m) molecule
(n) molecular formula
(o) neutron
(p) nucleus
(q) polyatomic
(r) radioactivity
(s) structural formula

True-False Statements

2.20 If a substance spontaneously emits alpha particles, we say it is *radioactive*.

2.21 The *atomic number* of a specific element is always constant.

2.22 The phrase "cobalt-60" means an isotope of cobalt that has a *mass number* of 60.

2.23 One *Angstrom* is equivalent to 10^{-8} m.

2.24 Elements in the *periodic table* are sequentially arranged in order of increasing mass number.

2.25 One *electronic charge* is equivalent to one coulomb.

2.26 He, Ne, and Ar are all members of a *family* of elements.

2.27 An example of the *law of multiple proportions* is the existence of NO_2 and NO.

2.28 Hg is a *nonmetallic* element.

2.29 Sr is a *metallic* element.

2.30 Antimony is a *metalloid*.

2.31 To specify a *nuclide*, only its atomic number is needed.

2.32 A *structural formula* of a substance shows the simplest ratio of elements in the chemical formula.

2.33 An important aspect of a *ball-and-stick model* is the showing of approximate bond angles.

2.34 NO is an *ionic* substance.

2.35 KI contains a *cation*.

2.36 The charge of the *anion* in $CaCl_2$ is 1−.

2.37 An example of a *polyatomic ion* is the bromide ion in $AlBr_3$.

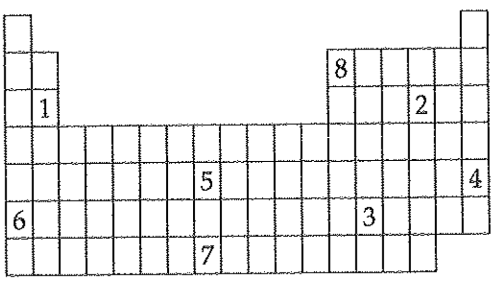
2.38 Methane, CH_4, is a *hydrocarbon*.

2.39 An *alkane* contains a carbon–carbon double bond.

2.40 Ethanol is an *alcohol* with two carbon atoms and an –OH group.

Problems and Short-Answer Questions

2.41 Which element in the following periodic table is most likely to form an anion?



(a) 1 (c) 3
(b) 2 (d) 4

2.42 Which element in the previous periodic table has the largest atomic weight?

(a) 5 (c) 7
(b) 6 (d) 8

2.43 John Dalton's model for the atom includes the postulate that all atoms of a given element are identical. Contrast this postulate to what we know about atoms today. Give at least one example to support your statements.

2.44 What are the three primary particles that comprise an atom of an element? How do they differ in charge and mass? What is the approximate range of diameters for atoms?

2.45 What does the symbol $^{201}_{80}X$ tell you about the isotope? From this information can you identify the element X?

2.46 Various columns in the periodic table are given names. Which groups are given the following names: Alkali metals, halogen, noble gases, and coinage metals?

2.47 What is a molecule? Can an element form a molecule or can molecules only form from combinations of different elements? What are seven important diatomic elements?

2.48 In the ensuing chapters you will often encounter ions. Does an atom become an ion by a change in the nuclear or electronic structure? Which is more likely to become a cation, potassium or sulfur? Which is more likely to become an anion, calcium or bromine?

2.49 What are the most likely charges of the elements Ca, P, and I? How does the periodic table help you answer this question?

2.50 If you are given a chemical formula such as FeO or ICl, is it possible to predict whether the substance is ionic or molecular? Explain.

2.51 Critique the following statements: The name of $FeCl_3$ is iron chloride. The name of NO_2 is nitrogen dioxygen.

2.52 Name the following anions: H^-; Cl^-; CN^-; $C_2H_3O_2^-$; and CrO_4^{2-}.

2.53 What is an acid? What clue in a chemical formula may help you identify a substance as an acid? When naming the binary acid HCl we use the prefix *hydro-* and add the *-ic* ending to the nonmetal to form hydrochloric acid. Why can't we simply name it chloric acid?

2.54 The name of a particular alcohol is hexanol. Is there sufficient information to determine both the number of carbon atoms and to which carbon atom the alcohol functional group is attached? If not, what additional information is needed?

2.55 The prefix cyclo- in the name of an alkane tells us that the carbon atoms form a ring. Is there sufficient information given to conclude that the alkane cyclopentane has ten hydrogen atoms? If not, what additional information is needed?

2.56 The chlorine (Cl) atom occurs in combination with many other elements.

(a) A $^{37}_{17}Cl^-$ nuclide contains how many protons, electrons, and neutrons?
(b) What is chlorine's elemental form?
(c) In the elemental state, chlorine is a green gas at room temperature. Is this color a physical or chemical property?
(d) Write the formula of a compound containing Cl^- with Cr^{3+}.
(e) 1 g of KCl is placed in 100 mL of water, and it dissolves with stirring. Identify the combination as one or more of the following: element, compound, heterogeneous mixture, homogeneous mixture, or solution.

2.57 Identify the following elements if an atom of each has the following:
(a) A 1+ charge and 11 protons;
(b) a 2+ charge and 36 electrons;
(c) a 2^- charge and an atomic number of 34.

2.58 Bromine has two naturally occurring isotopes: ^{79}Br (78.918 amu) and ^{81}Br (80.916 amu). The atomic weight of Br is 79.904 amu. What are the fractional abundances of ^{79}Br and ^{81}Br?

2.59 Complete the following table:

Symbol	Number of protons	Number of neutrons	Number of electrons	Charge
$^{90}_{38}Sr^{2+}$				
	92	143		0
		10	10	1−
	46	60		2+

2.60 Predict the chemical formulas of the compounds most likely to form when the following combine and indicate if the predicted substances are ionic or molecular: (a) Al and nitrate ion; (b) Mg and phosphate ion; (c) ammonium ion and Br.

2.61 Name the following compounds:
(a) $CaSO_4$ (f) H_2SO_4
(b) PF_5 (g) CO_2
(c) KBr (h) $HClO_4$
(d) $KHSO_4$ (i) $NaClO_3$
(e) Na_2S (j) $Cu(CN)_2$

2.62 Using the periodic table, determine which ions are not likely:
(a) Cl^+ (c) S^{2-}
(b) Cs^+ (d) Rb^-

2.63 Write the formula for each of the following:
(a) tin(IV) chloride; (d) dinitrogen trioxide;
(b) chromium(III) hydroxide; (e) cobalt(III) oxide;
(c) cesium cyanide; (f) calcium phosphate;
(g) osmium tetraoxide; (i) hypobromous acid;
(h) mercury(II) bromide; (j) hydroselenic acid.

2.64 Write the chemical formula for or give the name of each of the following oxyanions and oxyacids:
(a) HClO (e) $HMnO_4$
(b) ClO^- (f) MnO_4^-
(c) perchloric acid (g) H_2SO_3
(d) perchlorate ion (h) SO_3^{2-}

2.65 Write the chemical formulas for:
(a) ethane (c) butanol
(b) propane (d) methanol

Multiple-Choice Questions

2.66 Which experimental observation evidence indicates that cathode rays consist of electrons?
(a) Their path is curved in a magnetic field;
(b) Gravity significantly affects their path over a short distance;
(c) They have no effect on an electroscope;
(d) They are assigned a negative charge;
(e) All of the above are indications.

2.67 Which particle helped to explain discrepancies in atomic weights observed by early nineteenth century chemists?
(a) neutron (d) atom
(b) electron (e) molecule
(c) proton

2.68 The development of modern atomic theory took a leap forward with the proof that an atom is mostly empty space. Which of the following experiments was responsible for this proof?
(a) Millikan's oil drop
(b) cathode-ray deflection in a magnetic field
(c) the bombardment of gold foil with alpha particles
(d) separation of gases by gas chromatography
(e) the Wilson cloud chamber.

2.69 Which of the following cations are *correctly* named: (1) Sn^{2+}—tin(II) ion; (2) NH_4^+—ammonium ion; (3) Ca^{2+}—calcium ion?
(a) 1 and 2 (d) all of them
(b) 2 and 3 (e) none of them
(c) 1 and 3

2.70 Which of the following anions are *incorrectly* named: (1) IO^-—hyperiodate ion; (2) SO_3^{2-}—sulfite ion; (3) ClO_3^-—chlorate ion; (4) Cl^-—chloride ion; (5) NO_2^-—nitrate ion?
(a) 1 and 2 (d) all of them
(b) 1 and 3 (e) none of them
(c) 1 and 5

2.71 Which of the following compounds are *incorrectly* named: (1) $FeCl_3$—ferric chloride; (2) $HgCl_2$—mercurous chloride; (3) SnS_2—tin(IV) sulfide; (4) $KClO_2$—potassium chlorate?

(a) 1 and 2
(b) 2 and 4
(c) 3 and 4
(d) all of them
(e) none of them

2.72 Which of the following acids in water are *correctly* named: (1) H_2S—hydrosulfuric; (2) $HClO_4$—perchloric; (3) H_3PO_3—phosphorous; (4) HNO_2—nitrous; (5) H_2CO_3—carbonic?

(a) 1 and 2
(b) 2, 3, and 4
(c) 3, 4, and 5
(d) all of them
(e) none of them

2.73 Which of the following is the formula of phosphoric acid?

(a) H_2CO_3
(b) H_3P
(c) H_3PO_4
(d) H_3PO_2
(e) H_2PO_3

2.74 Which of the following is the most likely formula of the compound formed between Sr and S?

(a) SrS
(b) Sr_2S
(c) SrS_2
(d) Sr_2S_3
(e) Sr_3S_2

2.75 Which pair represents the correct chemical formulas for iron(II) oxide and iron(III) oxide?

(a) FeO and Fe_3O_2
(b) Fe_2O and Fe_2O_3
(c) FeO and Fe_2O_3
(d) FeO_2 and FeO
(e) FeO_2 and FeO_2

2.76 The oxyanions of the 3rd period have at most four oxygen atoms. Which of the following represents the correct chemical formulas for the perchlorate, sulfate, phosphate, and silicate ions? The silicate ion is not given in Chapter 2, but you can predict it by analogy.

(a) ClO^-, SO_4^{2-}, PO_4^{3-}, SiO_3^{3-}
(b) ClO_4^-, SO_4^{2-}, PO_4^{3-}, SiO_3^{3-}
(c) ClO_3^-, SO_4^{2-}, PO_4^{3-}, SiO_4^{3-}
(d) ClO_4^-, SO_4^{2-}, PO_4^{3-}, SiO_3^{4-}
(e) ClO_4^-, SO_4^{2-}, PO_4^{3-}, SiO_4^{4-}

2.77 What do these elements have in common: Na, K, and Cs?

1. metallic
2. nonmetallic
3. form cations
4. form anions
5. alkali metals
6. alkaline-earth metals
7. halogens

(a) 1 and 3
(b) 2 and 4
(c) 1, 4, and 7
(d) 1, 3, and 5
(e) 2, 4, and 6

2.78 What is the charge of the metal ion in $Co_3(PO_4)_2$?

(a) 1+
(b) 2+
(c) 3+
(d) 2−
(e) 3−

2.79 The use of mercury salts has been minimized in general chemistry laboratories because of their potential toxicity and damage to the environment if not disposed of properly. Which of the following represents the correct chemical formulas for mercury(I) chloride and mercury(II) sulfide?

(a) HgCl and HgS_2
(b) $HgCl_2$ and HgS_2
(c) Hg_2Cl and HgS_2
(d) Hg_2Cl_2 and HgS
(e) Hg_2Cl_2 and Hg_2S

2.80 Which is not an alkane?

(a) C_3H_8
(b) CH_4
(c) C_4H_{10}
(d) C_5H_8
(e) C_8H_{18}

SELF-TEST SOLUTIONS

2.1 (e). **2.2** (b).**2.3** (h). **2.4** (o). **2.5** (a). **2.6** (k). **2.7** (l).**2.8** (c). **2.9** (g). **2.10** (p). **2.11** (q). **2.12** (i). **2.13** (s). **2.14** (j). **2.15** (n). **2.16** (f). **2.17** (m). **2.18** (r). **2.19** (d). **2.20** True. **2.21** True. **2.22** True. **2.23** False. It equals 10^{-10} m. **2.24** False. In order of increasing atomic number. **2.25** False. 1.602×10^{-19} C = 1 electronic charge. **2.26** True. **2.27** True. **2.28** False. A metal. **2.29** True. **2.30** True. **2.31** False. Both its atomic number and mass number must be specified. **2.32** False. A structural formula is a picture that shows how atoms are attached within a molecule and it normally shows the actual number of elements, a molecular formula. **2.33** True. **2.34** False. It is a covalent substance. An ionic substance contains a metal and a nonmetal. **2.35** True. A cation is an ion with a positive charge, K^+. **2.36** True. **2.37** False. A polyatomic ion consists of atoms joined together as in a molecule but it contains a charge. The bromide ion consists of only one element; a polyatomic ion will have two or more different elements. **2.38** True. **2.39** False. An alkane contains all single carbon-carbon bonds. **2.40** True. **2.41** (b) Element "2," sulfur, forms the sulfide ion. Element "4" is a noble gas and does not readily form ions. **2.42** (c) Element "7" is Hs, hassium.

2.43 John Dalton believed that atoms were the basic building blocks of all matter and that all atoms of a particular element were identical. This concept survived until the discovery of isotopes. Isotopes are atoms of the same element that differ in the number of neutrons. For example, the element uranium has three naturally occurring isotopes: $^{233}_{92}U$, $^{235}_{92}U$, $^{238}_{92}U$.

2.44 The three primary particles are the proton, electron, and neutron. A proton has a positive charge and a mass of 1.0073 amu; an electron has a negative charge and a mass of 5.486×10^{-4} amu; and a neutron has no charge and a mass of 1.0087 amu. Most atoms have diameters between 1×10^{-10} m and 5×10^{-10} m.

2.45 The subscript tells you the atomic number of the element. In this case, the atomic number 80 uniquely identifies the element mercury. The superscript tells you the sum of protons and neutrons, 201. The difference, 201–80, indicates that the atom has 121 neutrons.

2.46 Refer to Table 2.3 in the text for a complete collection of names. The alkali metals form group 1A, the halogens form group 7A, the noble gases form group 8A, and the coinage metals—copper, silver, and gold—belong to group 1B.

2.47 A molecule is a substance that contains two or more atoms tightly bound together and these combined atoms act as a unit with its own unique properties. In the text you encounter six important elements that form diatomic molecules: H_2; O_2; N_2; F_2; Cl_2; Br_2; and I_2. A molecule can be composed of identical elements if they are combined and form a discrete unit.

2.48 An atom becomes an ion by the addition of electrons to form an anion or the removal of electrons to become a cation. A change in the number of protons in the nucleus would alter the identity of the element to a different one. Cations are generally formed by metals; thus, potassium is more likely to form a cation. Anions are generally formed by nonmetals; thus, bromine is more likely to become an anion.

2.49 Ca^{2+}, P^{3-}, and I^-. For an active metal, groups 1(1A) and 2(2A), the most likely positive charge is the number of the group in which it resides. For a nonmetal the most likely negative charge is the difference between its group number and the number of the last family of elements in the periodic table.

2.50 In Chapter 2 you are given some general statements about ionic and molecular compounds that help you predict the type. Ionic compounds are generally formed from combinations of metals and nonmetals, whereas molecular compounds are generally composed of nonmetals only. If you do this analysis, then FeO should be ionic and ICl should be molecular. Some substances do not follow these observations. For example, $BeCl_2$ has a significant amount of molecular character.

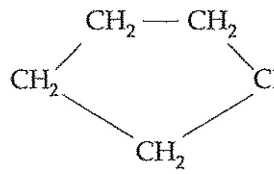
2.51 Iron is a transition element and exhibits more than one charge state. For elements of this type we indicate the charge as a Roman numeral, in parentheses, after the name of the element. Thus, $FeCl_3$ is named as iron(III) chloride. NO_2 is a molecular compound and for molecular compounds we name the first element as it is named in the periodic table. The second element is named as if it were a negative ion. For both elements the number of atoms is indicated using the appropriate prefix. NO_2 is named as nitrogen dioxide.

2.52 Hydride ion; chloride ion; cyanide ion; acetate ion; and chromate ion.

2.53 An acid is a hydrogen-containing compound that can form hydrogen ions (H^+) when dissolved in water. Generally, a compound is identified as an acid if it begins with a hydrogen element such as HCl, HNO_3, or $HC_2H_3O_2$. The name chloric acid refers to an oxyacid with the formula $HClO_3$. The prefix *hydro-* is used with binary acids to distinguish them from oxyacids.

2.54 There is sufficient information to determine hexanol has six carbon atoms. The prefix *hex-* means six carbon atoms in a hydrocarbon. However, there is insufficient information to determine the location of the –OH group. It could be at the end of the carbon chain but it might also be attached to an interior carbon atom with the –OH group. Additional information is needed in the name to identify the location of the carbon atom. In Chapter 2.25 additional nomenclature rules are provided to locate the carbon atom attached to the –OH group.

2.55 There is sufficient information. The prefix *pent-* in pentane tells us that it contains five carbon atoms. If we form a structural formula with five carbon atoms in a cyclic chain and then add hydrogen atoms so each carbon atom has four bonds, we obtain



Cyclopentane

From the structural picture we count ten hydrogen atoms.

2.56 (a) $^{37}_{17}Cl^-$ has 18 electrons (17 electrons for a neutral $^{37}_{17}Cl$ and one for the negative charge), 17 protons, and 20 neutrons.

(b) Cl_2 (similar to F_2, Br_2, and I_2).

(c) Physical. Its color does not determine its chemical behavior.

(d) $CrCl_3$. Three Cl^- ions are needed to balance the 3+ charge of Cr^{3+}.

(e) It is a homogeneous mixture, that is, a solution.

2.57 (a) The atomic number of the unknown element equals its number of protons: 11. The charge results from a loss of electrons, not the addition of protons. Na has an atomic number of 11.

(b) The 2+ charge results from the loss of two electrons. Thus in the neutral atom there are $36 + 2 = 38$ electrons and 38 protons. It has the atomic number 38, which is the atomic number of Sr.

(c) Se has an atomic number of 34.

2.58 Let x and y equal the fractional abundances of ^{79}Br and ^{81}Br, respectively. Then $x + y = 1$, as the sum of the fractional abundances must equal unity.

The fractional abundances, x and y, are solved for as follows:

$$AW(Br) = x(\text{amu of}^{79}\text{ Br}) + y(\text{amu of}^{81}\text{ Br})$$

$$79.904 \text{ amu} = x(78.918 \text{ amu}) + y(80.916 \text{ amu})$$

From the relationship $x + y = 1$ we write: $y = 1 - x$

$$79.904 \text{ amu} = x(78.918 \text{ amu}) + (1 - x)(80.916 \text{ amu})$$

Solving for x gives: 0.50651; and thus

$y = 1 - 0.50651 = 0.49349$

Thus elemental bromine consists of 50.651% ^{79}Br and 49.349% ^{81}Br.

2.59 This table is easily completed if you remember that the superscript number with a nuclide symbol equals the number of protons + the number of neutrons and that the subscript number equals the number of protons. The difference between the former and latter numbers equals the number of neutrons. Also, the number of electrons equals the number of protons adjusted for the charge of the nuclide.

Symbol	Number of protons	Number of neutrons	Number of electrons	Charge
$^{90}_{38}Sr^{2+}$	38	90 − 38 = 52	38 − 2 = 36	2+
$^{235}_{92}U$	92	143	92	0
$^{19}_{9}F^{-}$	9	10	9 + 1 = 10	1−
$^{106}_{46}Pd^{2+}$	46	60	46 − 2 = 44	2+

2.60 (a) $Al(NO_3)_3$.
(b) $Mg_3(PO_4)_2$.
(c) NH_4Br. They are all ionic substances. A molecular compound has all nonmetallic elements.

2.61 (a) calcium sulfate;
(b) phosphorus pentafluoride;
(c) potassium bromide;
(d) potassium hydrogen sulfate;
(e) sodium sulfide;
(f) sulfuric acid;
(g) carbon dioxide;
(h) perchloric acid;
(i) sodium chlorate;
(j) copper(II) cyanide.

2.62 (a) Chlorine is a nonmetal and thus is expected to form an ion with a *negative* charge;

(b) Cs^+ is correct because metals tend to form cations;

(c) S^{2-} is correct because nonmetals tend to form anions;

(d) Rb is a metal and thus is expected to form a cation, not an anion.

2.63 (a) $SnCl_4$;
(b) $Cr(OH)_3$;
(c) CsCN;
(d) N_2O_3;
(e) Co_2O_3;
(f) $Ca_3(PO_4)_2$;
(g) OsO_4;
(h) $HgBr_2$;
(i) HBrO;
(j) H_2Se.

2.64 (a) hypochlorous acid;
(b) hypochlorite ion;
(c) $HClO_4$;
(d) ClO_4^-;
(e) permanganic acid;
(f) permanganate ion;
(g) sulfurous acid;
(h) sulfite ion.

2.65 (a) C_2H_6
(b) C_3H_8
(c) C_4H_8OH
(d) CH_3OH

2.66 (a)

2.67 (a) Differing mixtures of isotopic nuclides (which differ in the number of neutrons) in samples of elements led to discrepancies in measurements of atomic weights.

2.68 (c). **2.69** (d). **2.70** (c). **2.71** (b). **2.72** (d). **2.73** (c). **2.74** (a).

2.75 (a) The oxide ion is O^{2-}. With iron(II), only one oxygen is needed to have the sum of charges equal zero. With iron(III), two iron(III) ions and three oxide ions are needed to have the sum of charges equal zero: $2(3+) + 3(2-) = 0$.

2.76 (e) All of the polyatomic ions in the question end in -ate. Thus, they possess the maximum number of oxygen atoms, four. The trend in their charges shows the effect of the first element being in different families.

2.77 (d)

2.78 (b) To determine the charge of the cobalt ion, you must first determine the charge of the polyatomic ion, the phosphate ion; it is 3−. Then apply the rule that the sum of charges for a substance must be zero. With three cobalt ions and two phosphate ions, the only way for this rule to apply is for the cobalt ion to be $2+ : 3(2+) = 0$.

2.79 (d) The mercury(I) ion is unusual in that it occurs in nature in a diatomic form.

2.80 (d) The general formula of a simple alkane is C_nH_{2n+2}. Only (d) does not obey this rule.