CHAPTER 2
ATOMS, MOLECULES, AND IONS

PROBLEM-SOLVING STRATEGIES AND TUTORIAL SOLUTIONS

TYPES OF PROBLEMS

Problem Type 1: Atomic number, Mass number, and Isotopes.

Problem Type 2: Empirical and Molecular Formulas.

Problem Type 3: Naming Compounds.
(a) Ionic compounds.
(b) Molecular compounds.
(c) Acids.
(d) Bases.

Problem Type 4: Formulas of Ionic Compounds.

PROBLEM TYPE 1: ATOMIC NUMBER, MASS NUMBER, AND ISOTOPES

The **atomic number (*Z*)** is the number of protons in the nucleus of each atom of an element.

EXAMPLE 2.1
What is the atomic number of an oxygen atom?

Solution: The atomic number is listed above each element in the periodic table. For oxygen, the atomic number is **8**, meaning that an oxygen atom has eight protons in the nucleus.

The **mass number (*A*)** is the total number of neutrons and protons present in the nucleus of an atom of an element.

$$\textbf{mass number} = \text{number of protons} + \text{number of neutrons}$$
$$\textbf{mass number} = \text{atomic number} + \text{number of neutrons}$$

EXAMPLE 2.2
A particular oxygen atom has nine neutrons in the nucleus. What is the mass number of this atom?

Strategy: Looking at a periodic table, you should find that every oxygen atom has an atomic number of 8. The number of neutrons is given, so we can solve for the mass number of this atom.

Solution:

$$\text{mass number} = \text{atomic number} + \text{number of neutrons}$$

$$\textbf{mass number} = 8 + 9 = \textbf{17}$$

Isotopes are atoms that have the same atomic number, but different mass numbers. For example, there are three isotopes of oxygen found in nature, oxygen-16, oxygen-17, and oxygen-18. The accepted way to denote the atomic number and mass number of an element X is as follows:

$$^{A}_{Z}\mathrm{X}$$

where, A = mass number
Z = atomic number

EXAMPLE 2.3
The three isotopes of oxygen found in nature are oxygen-16, -17, and -18. Write their isotopic symbols.

Strategy: The atomic number of oxygen is 8, so all isotopes of oxygen contain eight protons. The mass numbers are 16, 17, and 18, respectively.

Solution: $^{16}_{8}\mathrm{O}$ $^{17}_{8}\mathrm{O}$ $^{18}_{8}\mathrm{O}$

The number of **electrons** in an *atom* is equal to the number of protons.

number of electrons (atom) = number of protons = atomic number

The number of **electrons** in an *ion* is equal to the number of protons minus the charge on the ion.

number of electrons (ion) = number of protons – charge on the ion

EXAMPLE 2.4
What is the total number of fundamental particles (protons, neutrons, and electrons) in (a) an atom of $^{56}_{26}\mathrm{Fe}$ and (b) an $^{56}_{26}\mathrm{Fe}^{3+}$ ion?

Strategy: Both $^{56}_{26}\mathrm{Fe}$ and $^{56}_{26}\mathrm{Fe}^{3+}$ have the same atomic number and mass number, but the number of electrons will be different because one species is neutral and the other has a +3 charge.

Solution: For both (a) and (b):

number of protons = atomic number = **26**

and

number of neutrons = mass number – atomic number = 56 – 26 = **30**

However, the number of electrons for the above species differ.

(a) $^{56}_{26}\mathrm{Fe}$ is a neutral atom. Therefore,

number of electrons = number of protons = **26**

(b) $^{56}_{26}\mathrm{Fe}^{3+}$ is an ion with a +3 charge. Therefore,

number of electrons = number of protons – charge = 26 – (+3) = **23**

PRACTICE EXERCISE

1. How many protons, neutrons, and electrons are contained in each of the following atoms or ions?

(a) $^{19}\mathrm{F}$ **(b)** $^{79}\mathrm{Se}^{2-}$ **(c)** $^{40}\mathrm{Ca}$ **(d)** $^{48}\mathrm{Ti}^{4+}$

Text Problems: 2.14, 2.16, 2.18, 2.36

PROBLEM TYPE 2: EMPIRICAL AND MOLECULAR FORMULAS

A *molecular* formula shows the exact number of atoms of each element in the smallest unit of a substance. An *empirical formula* tells us which elements are present and the simplest whole-number ratio of their atoms. Empirical formulas are therefore the simplest chemical formulas; they are always written so that the subscripts in the molecular formulas are converted to the smallest possible whole numbers.

EXAMPLE 2.5
What is the empirical formula of each of the following compounds: (a) H_2O_2, (b) $C_6H_8O_6$, (c) $MgCl_2$, and (d) C_6H_6?

Strategy: An *empirical formula* tells us which elements are present and the *simplest* whole-number ratio of their atoms. Can you divide the subscripts in the formula by some factor to end up with smaller whole-number subscripts?

Solution:
(a) The simplest whole number ratio of the atoms in H_2O_2 is **HO**.

(b) The simplest whole number ratio of the atoms in $C_6H_8O_6$ is $\mathbf{C_3H_4O_3}$.

(c) The molecular formula as written contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

(d) The simplest whole number ratio of the atoms in C_6H_6 is **CH**.

PRACTICE EXERCISE

2. What is the empirical formula of each of the following compounds?
(a) C_2H_6O (b) $C_6H_{12}O_6$ (c) CH_3COOH (d) $C_{12}H_{22}O_{11}$

Text Problems: 2.46, 2.48

PROBLEM TYPE 3: NAMING COMPOUNDS

A. Naming ionic compounds

(1) **Metal cation has only one charge**. When naming ionic compounds, our reference for the names of cations and anions is Table 2.3 of the text. You should memorize the metal cations that have only one charge when they form ionic compounds. These include the alkali metals (Group 1A), which always have a +1 charge in ionic compounds, the alkaline earth metals (Group 2A), which always have a +2 charge in ionic compounds, and Al^{3+}, Ag^{+}, Cd^{2+}, and Zn^{2+}.

Since the metal cation has only one possible charge, we do not need to specify this charge in the compound. Therefore, the name of this type of ionic compound can be written simply by first naming the metal cation as it appears on the periodic table, followed by the nonmetallic anion. Anions from elements are named by changing the suffix to "-ide".

EXAMPLE 2.6
Name the following ionic compounds: (a) AlF_3, (b) Na_3N, (c) $Ba(NO_3)_2$.

Strategy: The metal cation in each of the compounds given has only one charge when forming ionic compounds. We do not specify this charge in naming the compound.

Solution:
(a) aluminum fluor*ide*
(b) sodium nitr*ide*
(c) barium nitrate

> **Tip:** There are a number of ions that contain more than one atom. These ions are called **polyatomic ions**. Nitrate, NO_3^-, is an example. Ask your instructor which polyatomic ions you should know.

(2) **Metals that form more than one type of cation.** Transition metals typically can form more than one type of cation when forming ionic compounds. If a metal can form cations of different charges, we need to use the Stock system. In the Stock system, Roman numerals are used to specify the charge of the cation. The Roman numeral (I) is used for one positive charge, (II) for two positive charges, (III) for three positive charges, and so on.

EXAMPLE 2.7
Name the following compounds: (a) FeO, (b) Fe_2O_3, (c) $HgSO_4$.

Strategy: The metals in the compounds above can form cations of different charges. We use the Stock system to name these compounds.

Solution:
(a) In this compound, the iron cation has a +2 charge, since oxide has a −2 charge (see Table 2.2 of the text). Therefore, the compound is named **iron(II) oxide**.

(b) In this compound, the iron cation has a +3 charge. Therefore, the compound is named **iron(III) oxide**.

(c) This compound contains the polyatomic ion sulfate, which has a −2 charge (SO_4^{2-}). Thus, the charge on mercury (Hg) is +2. The compound is named **mercury(II) sulfate**.

B. Naming molecular compounds

Unlike ionic compounds, molecular compounds contain discrete molecular units. They are usually composed of nonmetallic elements. There are two types of molecular compounds to consider.

(1) **Only one compound of the two elements exists.** If this is the case, you simply name the first element in the formula as it appears on the periodic table, followed by naming the second element with an "-ide" suffix.

EXAMPLE 2.8
Name the following compounds: (a) HF, (b) SiC.

(a) hydrogen fluoride
(b) silicon carbide

(2) **More than one compound composed of the two elements exists.** It is quite common for one pair of elements to form several different compounds. Therefore, we must be able to differentiate between the compounds. Greek prefixes are used to denote the number of atoms of each element present. See Table 2.4 in the text for the Greek prefixes used.

EXAMPLE 2.9
Name the following compounds: (a) CO, (b) CO_2, (c) N_2O_5, (d) SF_6.

(a) carbon *mono*xide
(b) carbon *di*oxide
(c) *di*nitrogen *pent*oxide
(d) sulfur *hexa*fluoride

> **Tip:** The prefix "mono-" may be omitted when naming the first element in a molecular compound (see carbon monoxide and carbon dioxide above). The absence of a prefix for the first element usually means that only one atom of that element is present in the molecule.

C. Naming acids

An acid is a substance that yields hydrogen ions (H^+) when dissolved in water. There are two types of acids to consider.

(1) **Acids that do not contain oxygen.** This type of acid contains one or more hydrogen atoms as well as an anionic group. To name these acids, you add the prefix "hydro-" to the anion name, change the "-ide" suffix of the anion to "-ic", and then add the word "acid" at the end.

EXAMPLE 2.10
Name the following binary acids: (a) HF(*aq*), (b) HCN(*aq*).

(a) *hydro*fluor*ic* acid
(b) CN^- is a polyatomic ion called cyanide. In the acid, the "-ide" suffix is changed to "-ic". The correct name is ***hydro*****cyan*****ic*** **acid**.

> **Tip:** The (*aq*) above means that the substance is dissolved in water. HF dissolved in water is an acid and is named hydrofluoric acid. However, HF in its pure state is a molecular compound and is named hydrogen fluoride.

(2) **Oxoacids.** This type of acid contains hydrogen, oxygen, and another element. To name oxoacids, you must look carefully at the anion name. If the suffix of the anion is "-ate", change the suffix to "-ic" and add the word "acid" at the end. If the suffix of the anion is "-ite", change the suffix to "-ous" and add the word "acid" at the end.

EXAMPLE 2.11
Name the following oxoacids: (a) HNO_2(*aq*), (b) $HClO_3$(*aq*).

(a) The NO_2^- polyatomic ion is called nitr*ite*. Simply change the suffix to "-ous" and add the word "acid". The correct name is **nitr*ous* acid**.
(b) The ClO_3^- polyatomic ion is called chlor*ate*. Simply change the suffix to "-ic" and add the work "acid". The correct name is **chlor*ic* acid**.

> **Tip:** If one O atom is added to the "-ic" acid, the acid is called "per...ic" acid. For example, $HClO_4$ is named perchloric acid. Compare this acid to chloric acid above. If one O atom is removed from the "-ous" acid, the acid is called "hypo...ous" acid. For example, HClO is named hypochlorous acid. Compare this to chlorous acid, $HClO_2$.

D. Naming Bases

A base is a substance that produces the hydroxide ion (OH^-) when dissolved in water. At this point, for naming purposes, we will only consider bases that contain the hydroxide ion. To name this type of base, simply name the metal cation first as it appears on the periodic table, then add "hydroxide".

EXAMPLE 2.12
Name the following bases: (a) KOH, (b) $Sr(OH)_2$.

(a) potassium hydroxide
(b) strontium hydroxide

PRACTICE EXERCISE

3. Name the following compounds:

   (a) $MgCl_2$ **(b)** $CuCl_2$ **(c)** HNO_3 **(d)** P_2O_5 **(e)** $Ca(OH)_2$

Text Problems: 2.58, 2.60

PROBLEM TYPE 4: FORMULAS OF IONIC COMPOUNDS

The formulas of ionic compounds are usually the same as their empirical formulas because ionic compounds do not consist of discrete molecular units. See Section 2.6 of the text if you need further information.

Ionic compounds are electrically neutral. In order for ionic compounds to be electrically neutral, the sum of the charges on the cation and anion in each formula unit must add up to zero. There are two possibilities to consider.

(1) If the charges on the cation and anion are numerically equal, no subscripts are necessary in the formula.
(2) If the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation.

You should memorize the metal cations that have only one charge when they form ionic compounds. These include the alkali metals (Group 1A), which always have a +1 charge in ionic compounds, the alkaline earth metals (Group 2A), which always have a +2 charge in ionic compounds, and Al^{3+}, Ag^{+}, Cd^{2+}, and Zn^{2+}. All other metals can have more than one possible positive charge when forming ionic compounds. This positive charge will be specified using a Roman numeral in the name of the compound.

You should also memorize the charges of polyatomic ions (see Table 2.3 of the text) and the common charges of monatomic anions based on their positions in the periodic table (see Table 2.2 of the text).

EXAMPLE 2.13
Write the formula for the ionic compound, magnesium oxide.

Strategy: The magnesium cation is an alkaline earth metal cation which always has a +2 charge in an ionic compound, and the oxide anion is −2 in an ionic compound (see Table 2.2 of the text).

Solution: Mg^{2+} and the oxide anion, O^{2-}, combine to form the ionic compound magnesium oxide. The sum of the charges is $+2 + (-2) = 0$, so no subscripts are necessary. The formula is **MgO**.

EXAMPLE 2.14
Write the formula for the ionic compound, iron(II) chloride.

Strategy: An iron cation can either have a +2 or +3 charge in an ionic compound. The Roman numeral, II, specifies that in this compound it is the +2 cation, Fe^{2+}. The chloride anion has a −1 charge in an ionic compound (see Table 2.2 of the text).

Solution: Fe^{2+} and the chloride anion, Cl^-, combine to form the ionic compound iron(II) chloride. The charges on the cation and anion are numerically different, so make the subscript of the cation (Fe^{2+}) numerically equal to the charge of the anion (subscript = 1). Also, make the subscript of the anion (Cl^-) numerically equal to the charge of the cation (subscript = 2). The formula is **$FeCl_2$**.

Tip: Check to make sure that the compound is electrically neutral by multiplying the charge of each ion by its subscript and then adding them together. The sum should equal zero.

$$\mathbf{(+2)(1) + (-1)(2) = 0}$$

PRACTICE EXERCISE

4. Write the correct formulas for the following ionic compounds:

   (a) Sodium oxide (b) Copper(II) nitrate (c) Aluminum oxide

Text Problems: 2.60 a, b, f, g, i, j

ANSWERS TO PRACTICE EXERCISES

1. (a) 9p, 10n, 9e
   (b) 34p, 45n, 36e
   (c) 20p, 20n, 20e
   (d) 22p, 26n, 18e

2. (a) C_2H_6O
   (b) CH_2O
   (c) CH_2O
   (d) $C_{12}H_{22}O_{11}$

3. (a) magnesium chloride
   (b) copper(II) chloride
   (c) nitric acid
   (d) diphosphorus pentoxide
   (e) calcium hydroxide

4. (a) Na_2O
   (b) $Cu(NO_3)_2$
   (c) Al_2O_3

SOLUTIONS TO SELECTED TEXT PROBLEMS

2.8 Note that you are given information to set up the unit factor relating meters and miles.

$$r_{\textbf{atom}} = 10^4 r_{\text{nucleus}} = 10^4 \times 2.0\ \cancel{\text{cm}} \times \frac{1\ \cancel{\text{m}}}{100\ \cancel{\text{cm}}} \times \frac{1\ \text{mi}}{1609\ \cancel{\text{m}}} = \mathbf{0.12\ mi}$$

2.14 Problem Type 1, Atomic number, Mass number, and Isotopes.

Strategy: The 239 in Pu-239 is the mass number. The **mass number (*A*)** is the total number of neutrons and protons present in the nucleus of an atom of an element. You can look up the atomic number (number of protons) on the periodic table.

Solution:

mass number = number of protons + number of neutrons

number of neutrons = mass number − number of protons = 239 − 94 = **145**

2.16

Isotope	$^{15}_{7}N$	$^{33}_{16}S$	$^{63}_{29}Cu$	$^{84}_{38}Sr$	$^{130}_{56}Ba$	$^{186}_{74}W$	$^{202}_{80}Hg$
No. Protons	7	16	29	38	56	74	80
No. Neutrons	8	17	34	46	74	112	122
No. Electrons	7	16	29	38	56	74	80

2.18 The accepted way to denote the atomic number and mass number of an element X is as follows:

$$^{A}_{Z}\mathbf{X}$$

where,

A = mass number
Z = atomic number

(a) $^{186}_{74}W$ **(b)** $^{201}_{80}Hg$

2.24 **(a)** Metallic character increases as you progress down a group of the periodic table. For example, moving down Group 4A, the nonmetal carbon is at the top and the metal lead is at the bottom of the group.

(b) Metallic character decreases from the left side of the table (where the metals are located) to the right side of the table (where the nonmetals are located).

2.26 F and Cl are Group 7A elements; they should have similar chemical properties. Na and K are both Group 1A elements; they should have similar chemical properties. P and N are both Group 5A elements; they should have similar chemical properties.

2.32 **(a)** This is a diatomic molecule that is a compound.
(b) This is a polyatomic molecule that is a compound.
(c) This is a polyatomic molecule that is the elemental form of the substance. It is not a compound.

2.34 There are more than two correct answers for each part of the problem.

(a) H_2 and F_2 **(b)** HCl and CO **(c)** S_8 and P_4
(d) H_2O and $C_{12}H_{22}O_{11}$ (sucrose)

2.36 The **atomic number (*Z*)** is the number of protons in the nucleus of each atom of an element. You can find this on a periodic table. The number of **electrons** in an *ion* is equal to the number of protons minus the charge on the ion.

number of electrons (ion) = number of protons – charge on the ion

Ion	K^+	Mg^{2+}	Fe^{3+}	Br^-	Mn^{2+}	C^{4-}	Cu^{2+}
No. protons	19	12	26	35	25	6	29
No. electrons	18	10	23	36	23	10	27

2.44 **(a)** The copper ion has a +1 charge and bromide has a –1 charge. The correct formula is CuBr.

(b) The manganese ion has a +3 charge and oxide has a –2 charge. The correct formula is Mn_2O_3.

(c) We have the Hg_2^{2+} ion and iodide (I^-). The correct formula is Hg_2I_2.

(d) Magnesium ion has a +2 charge and phosphate has a –3 charge. The correct formula is $Mg_3(PO_4)_2$.

2.46 Problem Type 2, Empirical and Molecular Formulas.

Strategy: An *empirical formula* tells us which elements are present and the *simplest* whole-number ratio of their atoms. Can you divide the subscripts in the formula by some factor to end up with smaller whole-number subscripts?

Solution:

(a) Dividing both subscripts by 2, the simplest whole number ratio of the atoms in Al_2Br_6 is $\mathbf{AlBr_3}$.

(b) Dividing all subscripts by 2, the simplest whole number ratio of the atoms in $Na_2S_2O_4$ is $\mathbf{NaSO_2}$.

(c) The molecular formula as written, $\mathbf{N_2O_5}$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

(d) The molecular formula as written, $\mathbf{K_2Cr_2O_7}$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

2.48 The molecular formula of ethanol is $\mathbf{C_2H_6O}$.

2.50 Compounds of metals with nonmetals are usually ionic. Nonmetal-nonmetal compounds are usually molecular.

Ionic: NaBr, BaF_2, CsCl.

Molecular: CH_4, CCl_4, ICl, NF_3

2.58 Problem Type 3, Naming Compounds.

Strategy: When naming ionic compounds, our reference for the names of cations and anions is Table 2.3 of the text. Keep in mind that if a metal can form cations of different charges, we need to use the Stock system. In the Stock system, Roman numerals are used to specify the charge of the cation. The metals that have only one charge in ionic compounds are the alkali metals (+1), the alkaline earth metals (+2), Ag^+, Zn^{2+}, Cd^{2+}, and Al^{3+}.

When naming acids, binary acids are named differently than oxoacids. For binary acids, the name is based on the nonmetal. For oxoacids, the name is based on the polyatomic anion. For more detail, see Section 2.7 of the text.

Solution:

(a) This is an ionic compound in which the metal cation (K^+) has only one charge. The correct name is **potassium hypochlorite**. Hypochlorite is a polyatomic ion with one less O atom than the chlorite ion, ClO_2^-

(b) **silver carbonate**

(c) This is an oxoacid that contains the nitrite ion, NO_2^-. The "-ite" suffix is changed to "-ous". The correct name is **nitrous acid.**

(d) **potassium permanganate** **(e)** **cesium chlorate** **(f)** **potassium ammonium sulfate**

(g) This is an ionic compound in which the metal can form more than one cation. Use a Roman numeral to specify the charge of the Fe ion. Since the oxide ion has a −2 charge, the Fe ion has a +2 charge. The correct name is **iron(II) oxide**.

(h) **iron(III) oxide**

(i) This is an ionic compound in which the metal can form more than one cation. Use a Roman numeral to specify the charge of the Ti ion. Since each of the four chloride ions has a −1 charge (total of −4), the Ti ion has a +4 charge. The correct name is **titanium(IV) chloride.**

(j) **sodium hydride** **(k)** **lithium nitride** **(l)** **sodium oxide**

(m) This is an ionic compound in which the metal cation (Na^+) has only one charge. The O_2^{2-} ion is called the peroxide ion. Each oxygen has a −1 charge. You can determine that each oxygen only has a −1 charge, because each of the two Na ions has a +1 charge. Compare this to sodium oxide in part (l). The correct name is **sodium peroxide.**

2.60 Problem Types 3 and 4.

Strategy: When writing formulas of molecular compounds, the prefixes specify the number of each type of atom in the compound.

When writing formulas of ionic compounds, the subscript of the cation is numerically equal to the charge of the anion, and the subscript of the anion is numerically equal to the charge on the cation. If the charges of the cation and anion are numerically equal, then no subscripts are necessary. Charges of common cations and anions are listed in Table 2.3 of the text. Keep in mind that Roman numerals specify the charge of the cation, *not* the number of metal atoms. Remember that a Roman numeral is not needed for some metal cations, because the charge is known. These metals are the alkali metals (+1), the alkaline earth metals (+2), Ag^+, Zn^{2+}, Cd^{2+}, and Al^{3+}.

When writing formulas of oxoacids, you must know the names and formulas of polyatomic anions (see Table 2.3 of the text).

Solution:

(a) The Roman numeral I tells you that the Cu cation has a +1 charge. Cyanide has a −1 charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is **CuCN**.

(b) Strontium is an alkaline earth metal. It only forms a +2 cation. The polyatomic ion chlorite, ClO_2^-, has a −1 charge. Since the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is **$Sr(ClO_2)_2$**.

(c) Perbromic tells you that the anion of this oxoacid is perbromate, BrO_4^-. The correct formula is **$HBrO_4$(*aq*)**. Remember that (*aq*) means that the substance is dissolved in water.

(d) Hydroiodic tells you that the anion of this binary acid is iodide, I^-. The correct formula is **HI(*aq*)**.

(e) Na is an alkali metal. It only forms a +1 cation. The polyatomic ion ammonium, NH_4^+, has a +1 charge and the polyatomic ion phosphate, PO_4^{3-}, has a −3 charge. To balance the charge, you need 2 Na^+ cations. The correct formula is $\mathbf{Na_2(NH_4)PO_4}$.

(f) The Roman numeral II tells you that the Pb cation has a +2 charge. The polyatomic ion carbonate, CO_3^{2-}, has a −2 charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is $\mathbf{PbCO_3}$.

(g) The Roman numeral II tells you that the Sn cation has a +2 charge. Fluoride has a −1 charge. Since the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is $\mathbf{SnF_2}$.

(h) This is a molecular compound. The Greek prefixes tell you the number of each type of atom in the molecule. The correct formula is $\mathbf{P_4S_{10}}$.

(i) The Roman numeral II tells you that the Hg cation has a +2 charge. Oxide has a −2 charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is **HgO**.

(j) The Roman numeral I tells you that the Hg cation has a +1 charge. However, this cation exists as Hg_2^{2+}. Iodide has a −1 charge. You need two iodide ion to balance the +2 charge of Hg_2^{2+}. The correct formula is $\mathbf{Hg_2I_2}$.

(k) This is a molecular compound. The Greek prefixes tell you the number of each type of atom in the molecule. The correct formula is $\mathbf{SeF_6}$.

2.62 Changing the electrical charge of an atom usually has a major effect on its chemical properties. The two electrically neutral carbon isotopes should have nearly identical chemical properties.

2.64 Atomic number = 127 − 74 = 53. This anion has 53 protons, so it is an iodide ion. Since there is one more electron than protons, the ion has a −1 charge. The correct symbol is $\mathbf{I^-}$.

2.66 NaCl is an ionic compound; it doesn't form molecules.

2.68 The species and their identification are as follows:

(a)	SO_2	molecule and compound	**(g)**	O_3	element and molecule
(b)	S_8	element and molecule	**(h)**	CH_4	molecule and compound
(c)	Cs	element	**(i)**	KBr	compound
(d)	N_2O_5	molecule and compound	**(j)**	S	element
(e)	O	element	**(k)**	P_4	element and molecule
(f)	O_2	element and molecule	**(l)**	LiF	compound

2.70 **(a)** Ammonium is NH_4^+, not NH_3^+. The formula should be $\mathbf{(NH_4)_2CO_3}$.

(b) Calcium has a +2 charge and hydroxide has a −1 charge. The formula should be $\mathbf{Ca(OH)_2}$.

(c) Sulfide is S^{2-}, not SO_3^{2-}. The correct formula is **CdS**.

(d) Dichromate is $Cr_2O_7^{2-}$, not $Cr_2O_4^{2-}$. The correct formula is $\mathbf{ZnCr_2O_7}$.

2.72 **(a)** Ionic compounds are typically formed between metallic and nonmetallic elements.

(b) In general the transition metals, the actinides and lanthanides have variable charges.

2.74 The symbol ^{23}Na provides more information than $_{11}Na$. The mass number plus the chemical symbol identifies a specific isotope of Na (sodium) while combining the atomic number with the chemical symbol tells you nothing new. Can other isotopes of sodium have different atomic numbers?

2.76 Mercury (Hg) and bromine (Br_2)

2.78 H_2, N_2, O_2, F_2, Cl_2, He, Ne, Ar, Kr, Xe, Rn

2.80 They do not have a strong tendency to form compounds. Helium, neon, and argon are chemically inert.

2.82 All isotopes of radium are radioactive. It is a radioactive decay product of uranium-238. Radium itself does *not* occur naturally on Earth.

2.84 Argentina is named after silver (argentum, Ag).

2.86 **(a)** NaH, sodium hydride **(b)** B_2O_3, diboron trioxide **(c)** Na_2S, sodium sulfide **(d)** AlF_3, aluminum fluoride **(e)** OF_2, oxygen difluoride **(f)** $SrCl_2$, strontium chloride

2.88 All of these are molecular compounds. We use prefixes to express the number of each atom in the molecule. The names are nitrogen trifluoride (NF_3), phosphorus pentabromide (PBr_5), and sulfur dichloride (SCl_2).

2.90

Cation	Anion	Formula	Name
$\mathbf{Mg^{2+}}$	$\mathbf{HCO_3^-}$	$\mathbf{Mg(HCO_3)_2}$	Magnesium bicarbonate
$\mathbf{Sr^{2+}}$	$\mathbf{Cl^-}$	$SrCl_2$	**Strontium chloride**
Fe^{3+}	NO_2^-	$\mathbf{Fe(NO_2)_3}$	**Iron(III) nitrite**
$\mathbf{Mn^{2+}}$	$\mathbf{ClO_3^-}$	$\mathbf{Mn(ClO_3)_2}$	Manganese(II) chlorate
$\mathbf{Sn^{4+}}$	$\mathbf{Br^-}$	$SnBr_4$	**Tin(IV) bromide**
Co^{2+}	PO_4^{3-}	$\mathbf{Co_3(PO_4)_2}$	**Cobalt(II) phosphate**
Hg_2^{2+}	I^-	$\mathbf{Hg_2I_2}$	**Mercury(I) iodide**
$\mathbf{Cu^+}$	$\mathbf{CO_3^{2-}}$	Cu_2CO_3	**Copper(I) carbonate**
$\mathbf{Li^+}$	$\mathbf{N^{3-}}$	$\mathbf{Li_3N}$	Lithium nitride
Al^{3+}	S^{2-}	$\mathbf{Al_2S_3}$	**Aluminum sulfide**

2.92 The change in energy is equal to the energy released. We call this ΔE. Similarly, Δm is the change in mass. Because $m = \dfrac{E}{c^2}$, we have

$$\Delta m = \frac{\Delta E}{c^2} = \frac{(1.715 \times 10^3\text{ kJ}) \times \dfrac{1000\text{ J}}{1\text{ kJ}}}{(3.00 \times 10^8\text{ m/s})^2} = 1.91 \times 10^{-11}\text{ kg} = \mathbf{1.91 \times 10^{-8}\text{ g}}$$

Note that we need to convert kJ to J so that we end up with units of kg for the mass. $\left(1\text{ J} = \dfrac{1\text{ kg}\cdot\text{m}^2}{\text{s}^2}\right)$

We can add together the masses of hydrogen and oxygen to calculate the mass of water that should be formed.

$$12.096\ \text{g} + 96.000 = 108.096\ \text{g}$$

The predicted change (loss) in mass is only 1.91×10^{-8} g which is too small a quantity to measure. Therefore, for all practical purposes, the law of conservation of mass is assumed to hold for ordinary chemical processes.

2.94 **(a)** Rutherford's experiment is described in detail in Section 2.2 of the text. From the average magnitude of scattering, Rutherford estimated the number of protons (based on electrostatic interactions) in the nucleus.

(b) Assuming that the nucleus is spherical, the volume of the nucleus is:

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(3.04 \times 10^{-13}\ \text{cm})^3 = 1.18 \times 10^{-37}\ \text{cm}^3$$

The density of the nucleus can now be calculated.

$$\boldsymbol{d} = \frac{m}{V} = \frac{3.82 \times 10^{-23}\ \text{g}}{1.18 \times 10^{-37}\ \text{cm}^3} = \mathbf{3.24 \times 10^{14}\ g/cm^3}$$

To calculate the density of the space occupied by the electrons, we need both the mass of 11 electrons, and the volume occupied by these electrons.

The mass of 11 electrons is:

$$11\ \text{electrons} \times \frac{9.1095 \times 10^{-28}\ \text{g}}{1\ \text{electron}} = 1.0020 \times 10^{-26}\ \text{g}$$

The volume occupied by the electrons will be the difference between the volume of the atom and the volume of the nucleus. The volume of the nucleus was calculated above. The volume of the atom is calculated as follows:

$$186\ \text{pm} \times \frac{1 \times 10^{-12}\ \text{m}}{1\ \text{pm}} \times \frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}} = 1.86 \times 10^{-8}\ \text{cm}$$

$$V_{\text{atom}} = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(1.86 \times 10^{-8}\ \text{cm})^3 = 2.70 \times 10^{-23}\ \text{cm}^3$$

$$V_{\text{electrons}} = V_{\text{atom}} - V_{\text{nucleus}} = (2.70 \times 10^{-23}\ \text{cm}^3) - (1.18 \times 10^{-37}\ \text{cm}^3) = 2.70 \times 10^{-23}\ \text{cm}^3$$

As you can see, the volume occupied by the nucleus is insignificant compared to the space occupied by the electrons.

The density of the space occupied by the electrons can now be calculated.

$$\boldsymbol{d} = \frac{m}{V} = \frac{1.0020 \times 10^{-26}\ \text{g}}{2.70 \times 10^{-23}\ \text{cm}^3} = \mathbf{3.71 \times 10^{-4}\ g/cm^3}$$

The above results do support Rutherford's model. Comparing the space occupied by the electrons to the volume of the nucleus, it is clear that most of the atom is empty space. Rutherford also proposed that the nucleus was a *dense* central core with most of the mass of the atom concentrated in it. Comparing the density of the nucleus with the density of the space occupied by the electrons also supports Rutherford's model.

2.96 **(a)**

Ethane	Acetylene
2.65 g C	4.56 g C
0.665 g H	0.383 g H

Let's compare the ratio of the hydrogen masses in the two compounds. To do this, we need to start with the same mass of carbon. If we were to start with 4.56 g of C in ethane, how much hydrogen would combine with 4.56 g of carbon?

$$0.665\text{ g H} \times \frac{4.56\text{ g C}}{2.65\text{ g C}} = 1.14\text{ g H}$$

We can calculate the ratio of H in the two compounds.

$$\frac{1.14\text{ g}}{0.383\text{ g}} \approx 3$$

This is consistent with the Law of Multiple Proportions which states that if two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers. In this case, the ratio of the masses of hydrogen in the two compounds is 3:1.

(b) For a given amount of carbon, there is 3 times the amount of hydrogen in ethane compared to acetylene. Reasonable formulas would be:

Ethane	Acetylene
CH_3	CH
C_2H_6	C_2H_2

2.98 The mass number is the sum of the number of protons and neutrons in the nucleus.

$$\text{Mass number} = \text{number of protons} + \text{number of neutrons}$$

Let the atomic number (number of protons) equal A. The number of neutrons will be $1.2A$. Plug into the above equation and solve for A.

$$55 = A + 1.2A$$
$$A = 25$$

The element with atomic number 25 is **manganese, Mn**.

2.100 The acids, from left to right, are chloric acid, nitrous acid, hydrocyanic acid, and sulfuric acid.