Burdge Chemistry 4e Chapter 2 ISM.pdf

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Chapter 2

Atoms, Molecules, and Ions

Practice Problems C

2.1 (ii) and (iii)

2.2 (i) 14N, (ii) 21Na, (iii) 15O

2.3 (i) 2.9177 g

(ii) 3.4679 g

(iii) 3.4988 g

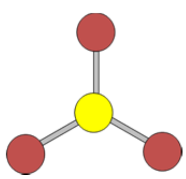
2.4 Fe(NO3)2, iron(II) nitrate

2.5 CuSO3, copper(II) sulfite

2.6 (i) and (iv)

2.7 4 ethanol molecules; all C used, 1 O left over, 3 H left over

2.8 selenium hexachloride

2.9 

2.10 (ii) and (iv)

2.11 formaldehyde and glucose

Applying What You've Learned

a) There are 54 – 26 = 28 neutrons in the ^{54}Fe nucleus, 30 neutrons in the ^{56}Fe nucleus, 31 neutrons in the ^{57}Fe nucleus, and 32 neutrons in the ^{58}Fe nucleus.

b) The average atomic mass of iron is 55.845 amu.

c) The molecular formula for ascorbic acid is $C_6H_8O_6$.

d) The empirical formula for ascorbic acid is $C_3H_4O_3$.

e) The formula for ferrous sulfate is $FeSO_4$.

Questions and Problems

2.1 1. Elements are composed of extremely small particles called atoms. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.

2. Compounds are composed of atoms of more than one element. In any given compound, the same types of atoms are always present in the same relative numbers.

3. A chemical reaction rearranges atoms in chemical compounds; it does not create or destroy them.

2.2 The law of definite proportions states that different samples of a given compound always contain the same elements in the same mass ratio.

The law of multiple proportions states that if two elements can combine to form more than one compound with each other, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

2.3

$$\frac{\text{ratio of S to O in compound 1}}{\text{ratio of S to O in compound 2}} = \frac{1.002}{0.668} \approx 1.5$$

2.4

$$\frac{\text{ratio of P to Cl in compound 1}}{\text{ratio of P to Cl in compound 2}} = \frac{0.2912}{0.1747} \approx \mathbf{1.667} \approx \mathbf{5:3}$$

2.5

$$\frac{\text{ratio of F to S in } S_2F_{10}}{\text{ratio of F to S in } SF_4} = \frac{2.962}{2.370} \approx 1.250$$

$$\frac{\text{ratio of F to S in } SF_6}{\text{ratio of F to S in } SF_4} = \frac{3.555}{2.370} \approx 1.5$$

$$\frac{\text{ratio of F to S in } SF_4}{\text{ratio of F to S in } SF_4} = 1$$

$$\text{ratio in } SF_6 : \text{ratio in } S_2F_{10} : \text{ratio in } SF_4 = 1.5 : 1.25 : 1$$

Multiply through to get all whole numbers. $4 \cdot (1.5 : 1.25 : 1) = \mathbf{6 : 5 : 4}$

2.6 $$\frac{\text{ratio of O to Fe in FeO}}{\text{ratio of O to Fe in Fe}_2\text{O}_3} = \frac{0.2865}{0.4297} \approx \mathbf{0.667} \approx \mathbf{2:3}$$

2.7 $$\frac{\text{g blue: 1.00 g red (right)}}{\text{g blue: 1.00 g red (left)}} = \frac{2/3}{1/1} \approx \mathbf{0.667} \approx \mathbf{2:3}$$

2.8 $$\frac{\text{g green: 1.00 g orange (right)}}{\text{g green: 1.00 g orange (left)}} = \frac{4/2}{3/1} \approx \mathbf{0.667:1} \approx \mathbf{2:3}$$

2.9 a. An α particle is a positively charged particle consisting of two protons and two neutrons, emitted in radioactive decay or nuclear fission.

b. A β particle is a high-speed electron, especially emitted in radioactive decay.

c. γ rays are high-energy electromagnetic radiation emitted by radioactive decay.

d. X-rays are a form of electromagnetic radiation similar to light but of shorter wavelength.

2.10 alpha rays, beta rays, and gamma rays

2.11 α particles are deflected away from positively charged plates. Cathode rays are drawn toward positively charged plates. Protons are positively charged particles in the nucleus. Neutrons are electrically neutral subatomic particles in the nucleus. Electrons are negatively charged particles that are distributed around the nucleus.

2.12 J .J. Thomson determined the ratio of electric charge to the mass of an individual electron.

R. A. Millikan calculated the mass of an individual electron and proved the charge on each electron was exactly the same.

Ernest Rutherford proposed that an atom's positive charges are concentrated in the nucleus and that most of the atom is empty space.

James Chadwick discovered neutrons.

2.13 Rutherford bombarded gold foil with α particles. Most of them passed through the foil, while a small proportion were deflected or reflected. Thus, most of the atom must be empty space through which the α particles could pass without encountering any obstructions.

2.14 First, convert 1 cm to picometers.

$$1 \text{ cm} \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{1 \text{ pm}}{1 \times 10^{-12} \text{ m}} = 1 \times 10^{10} \text{ pm}$$

$$(1 \times 10^{10} \text{ pm}) \times \frac{1 \text{ He atom}}{1 \times 10^{2} \text{ pm}} = \mathbf{1 \times 10^{8} \text{ He atoms}}$$

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

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

2.16 Atomic number is the number of protons in the nucleus of each atom of an element. It determines the chemical identity of the element. There are 2 protons in each atom of helium-4.

Mass number is the total number of neutrons and protons present in the nucleus of an atom of an element. The mass number of helium-4 is 4. There are (4 – 2) = 2 neutrons in each atom.

Because atoms are electrically neutral, the number of protons and electrons must be equal. The atomic number is also the number of electrons in each atom.

2.17 The atomic number is the number of protons in the nucleus. It determines the chemical identity of the element. If an atom has a different number of protons (a different atomic number), it is a different element.

2.18 isotopes

2.19 *X* is the element symbol. It indicates the chemical identity of the atom.

A is the mass number. It is the number of protons plus the number of neutrons.

Z is the atomic number. It is the number of protons.

2.20 For iron, the atomic number *Z* is 26. Therefore the mass number *A* is:

$$\boldsymbol{A} = 26 + 28 = \mathbf{54}$$

2.21 **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**

2.22

Isotope	${}^{3}_{2}He$	${}^{4}_{2}He$	${}^{24}_{12}Mg$	${}^{25}_{12}Mg$	${}^{48}_{22}Ti$	${}^{79}_{35}Br$	${}^{195}_{78}Pt$
No. Protons	2	2	12	12	22	35	78
No. Neutrons	1	2	12	13	26	44	117

2.23

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. Electrons	7	16	29	38	56	74	80
No. Neutrons	8	17	34	46	74	112	122

2.24 a. ${}^{23}_{11}Na$ b. ${}^{64}_{28}Ni$ c. ${}^{115}_{50}Sn$ d. ${}^{42}_{20}Ca$

2.25 The accepted way to denote the atomic number and mass number of an element X is ${}^{A}_{Z}X$ where ***A*** = mass number and ***Z*** = atomic number.

a. ${}^{186}_{74}W$ b. ${}^{201}_{80}Hg$ c. ${}^{76}_{34}Se$ d. ${}^{239}_{94}Pu$

2.26 a. 11 b. 25 c. 80 d. 199

2.27 a. 20 b. 32 c. 78 d. 198

2.28 ${}^{198}Au$: 119 neutrons, ${}^{47}Ca$: 27 neutrons, ${}^{60}Co$: 33 neutrons, ${}^{18}F$: 9 neutrons, ${}^{125}I$: 72 neutrons, ${}^{131}I$: 78 neutrons, ${}^{42}K$: 23 neutrons, ${}^{43}K$: 24 neutrons, ${}^{24}Na$: 13 neutrons, ${}^{32}P$: 17 neutrons, ${}^{85}Sr$: 47 neutrons, ${}^{99}Tc$: 56 neutrons.

2.29 The periodic table is a chart in which elements having similar chemical and physical properties are grouped together.

2.30 A metal is a good conductor of heat and electricity, whereas a nonmetal is usually a poor conductor of heat and electricity.

2.31 Answers will vary.

2.32 Answers will vary.

2.33 Strontium has similar chemical properties to calcium, which is an important mineral for humans.

2.34 Helium and Selenium are nonmetals whose name ends with *ium*. (Tellurium is a metalloid whose name ends in *ium*.)

2.35 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).

2.36 a. Li (0.53 g/cm^3) K (0.86 g/cm^3) H_2O (0.98 g/cm^3)

b. Au (19.3 g/cm^3) Pt (21.4 g/cm^3) Hg (13.6 g/cm^3)

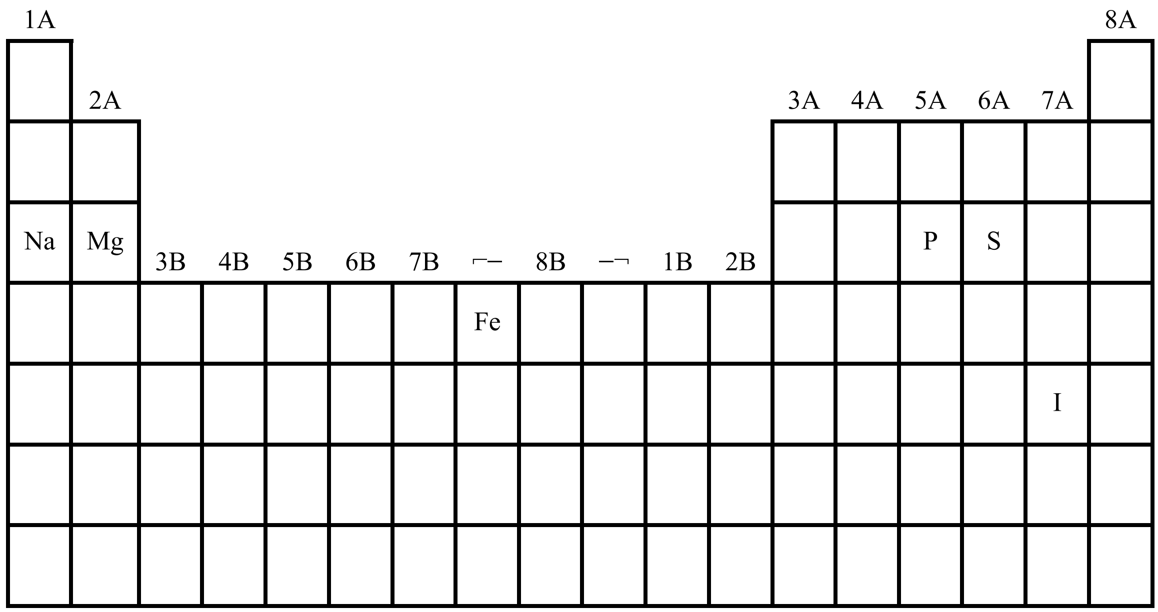
c. Os (22.6 g/cm^3)

d. Te (6.24 g/cm^3)

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

2.38 I and Br (both in Group 7A), O and S (both in Group 6A), Ca and Ba (both in Group 2A)

2.39



Atomic number 26, iron, Fe, (present in hemoglobin for transporting oxygen)

Atomic number 53, iodine, I, (present in the thyroid gland)

Atomic number 11, sodium, Na, (present in intra- and extra-cellular fluids)

Atomic number 15, phosphorus, P, (present in bones and teeth)

Atomic number 16, sulfur, S, (present in proteins)

Atomic number 12, magnesium, Mg, (present in chlorophyll molecules)

2.40 An atomic mass unit (amu) is defined as a mass exactly equal to one-twelfth the mass of one carbon-12 atom. Setting the atomic mass of carbon-12 at 12 amu provides the standard for measuring the atomic mass of the other elements, since the mass of a single atom cannot be measured.

2.41 The mass of a carbon-12 atom is exactly 12 amu. The mass on the periodic table is the average mass of naturally occurring carbon, which is a mixture of several carbon isotopes.

2.42 The average mass of the naturally occurring isotopes of gold, taking into account their natural abundances, is 197.0 amu.

2.43 To calculate the average atomic mass of an element, you must know the identity and natural abundances of all naturally occurring isotopes of the element.

2.44 $$(78.9183361 \text{ amu})(0.5069) + (80.916289 \text{ amu})(0.4931) = \mathbf{79.90\ amu}$$

2.45 (203.973020 amu)(0.014) + (205.974440 amu)(0.241) +(206.975872 amu)(0.221) + (207.976627 amu)(0.524) = **207.2 amu**

2.46 The fractional abundances of the two isotopes of Tl must add to 1. Therefore, we can write

$$(202.972320 \text{ amu})(x) + (204.974401 \text{ amu})(1 - x) = 204.4 \text{ amu}$$

Solving for x gives 0.2869. Therefore, the natural abundances of ^{203}Tl and ^{205}Tl are **28.69%** and **71.31%**, respectively.

2.47 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (not percent) will give the contribution to the average atomic mass of that particular isotope.

It would seem that there are two unknowns in this problem, the fractional abundance of ^{6}Li and the fractional abundance of ^{7}Li. However, these two quantities are not independent of each other; they are related by the fact that they must sum to 1. Start by letting x be the fractional abundance of ^{6}Li. Since the sum of the two fractional abundances must be 1, we can write

$$(6.0151 \text{ amu})(x) + (7.0160 \text{ amu})(1 - x) = 6.941 \text{ amu}$$

Solution: Solving for x gives 0.075, which corresponds to the fractional abundance of ^{6}Li. The fractional abundance of ^{7}Li is $(1 - x) = 0.925$. Therefore, the natural abundances of ^{6}Li and ^{7}Li are **7.5%** and **92.5%**, respectively.

2.48 The conversion factor required is $\dfrac{6.022 \times 10^{23} \text{ amu}}{1 \text{ g}}$

$$13.2 \text{ amu} \times \frac{1 \text{ g}}{6.022 \times 10^{23} \text{ amu}} = \mathbf{2.19 \times 10^{-23}\ g}$$

2.49 The conversion factor required is $\dfrac{6.022 \times 10^{23} \text{ amu}}{1 \text{ g}}$

$$8.4 \text{ g} \times \frac{6.022 \times 10^{23} \text{ amu}}{1 \text{ g}} = \mathbf{5.1 \times 10^{24}\ amu}$$

2.50 Answers will vary.

2.51 An ionic compound consists of anions and cations. The ratio of anions and cations is such that the net charge is zero.

2.52 The formulas of ionic compounds are generally empirical formulas because an ionic compound consists of a vast array of interspersed cations and anions called a lattice, not discrete molecular units.

2.53 The Stock system uses Roman numerals to indicate the charge on cations of metals that commonly have more than one possible charge. This eliminates the need to know which charges are common on all the transition metals.

2.54 Answers will vary. Example: NH_4Cl

2.55 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	Na^+	Ca^{2+}	Al^{3+}	Fe^{2+}	I^-	F^-	S^{2-}	O^{2-}	N^{3-}
No. protons	11	20	13	26	53	9	16	8	7
No. electrons	10	18	10	24	54	10	18	10	10

2.56

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

2.57 a. Sodium ion has a +1 charge and oxide has a –2 charge. The correct formula is **Na_2O**.

b. The iron ion has a +2 charge and sulfide has a –2 charge. The correct formula is **FeS**.

c. The correct formula is **$Co_2(SO_4)_3$**.

d. Barium ion has a +2 charge and fluoride has a –1 charge. The correct formula is **BaF_2**.

2.58 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 **$Mg_3(PO_4)_2$**.

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

Ionic: LiF, $BaCl_2$, KCl

Molecular: $SiCl_4$, B_2H_6, C_2H_4

2.60 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

2.61 **Strategy:** When naming ionic compounds, our reference for the names of cations and anions are Tables 2.8 and 2.9 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 dihydrogen phosphate**.

b. This is an ionic compound in which the metal cation (K^+) has only one charge. The correct name is **potassium hydrogen phosphate**

c. This is molecular compound. In the gas phase, the correct name is **hydrogen bromide**.

d. The correct name of this compound in water is **hydrobromic acid**.

e. This is an ionic compound in which the metal cation (Li^+) has only one charge. The correct name is **lithium carbonate**.

f. This is an ionic compound in which the metal cation (K^+) has only one charge. The correct name is **potassium dichromate**.

g. This is an ionic compound in which the cation is a polyatomic ion with a charge of +1. The anion is an oxoanion with one less O atom than the corresponding –ate ion (nitrate). The correct name is **ammonium nitrite**.

h. The oxoanion in this acid is analogous to the chlorate ion. The correct name of this compound is **hydrogen iodate (in water, iodic acid)**

i. This is a molecular compound. We use a prefix to denote how many F atoms it contains. The correct name is **phosphorus pentafluoride**.

j. This is a molecular compound. We use prefixes to denote the numbers of both types of atom. The correct name is **tetraphosphorus hexoxide**.

k. This is an ionic compound in which the metal cation (Cd^{2+}) has only one charge. The correct name is **cadmium iodide**.

l. This is an ionic compound in which the metal cation (Sr^{2+}) has only one charge. The correct name is **strontium sulfate**.

m. This is an ionic compound in which the metal cation (Al^{3+}) has only one charge. The correct name is **aluminum hydroxide**.

2.62 a. potassium hypochlorite

b. silver carbonate

c. nitrous acid

h. iron(III) oxide

i. titanium(IV) chloride

j. sodium hydride

d. potassium permanganate

e. cesium chlorate

f. potassium ammonium sulfate

g. iron(II) oxide

k. lithium nitride

l. sodium oxide

m. sodium peroxide

2.63 **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 Tables 2.8 and 2.9 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 2.9 of the text).

Solution: a. Rubidium is an alkali metal. It only forms a +1 cation. The polyatomic ion nitrite, NO_2^-, has a −1 charge. Because the charges on the cation and anion are numerically equal, the ions combine in a one-to-one ratio. The correct formula is **$RbNO_2$**.

b. Potassium is an alkali metal. It only forms a +1 cation. The anion, sulfide, has a charge of −2. Because 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 **K_2S**.

c. Sodium is an alkali metal. It only forms a +1 cation. The anion is the *hydrogen sulfide* ion (the sulfide ion plus one hydrogen), HS^-. Because the charges are numerically the same, the ions combine in a one-to-one ratio. The correct formula is **NaHS**.

d. Magnesium is an alkaline earth metal. It only forms a +2 cation. The polyatomic phosphate anion has a charge of −3, PO_4^{3-}. Because 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 **$Mg_3(PO_4)_2$**. Note that for its subscript to be changed, a polyatomic ion must be enclosed in parentheses.

e. Calcium is an alkaline earth metal. It only forms a +2 cation. The polyatomic ion hydrogen phosphate, HPO_4^{2-}, has a −2 charge. Because the charges are numerically the same, the ions combine in a one-to-one ratio. The correct formula is **$CaHPO_4$**.

f. Lead (II), Pb^{2+}, is a cation with a charge +2. The polyatomic ion carbonate, CO_3^{2-}, has a −2 charge. Because the charges on the cation and anion are numerically equal, the ions combine in a one-to-one ratio. The correct formula is **$PbCO_3$**.

g. Tin (II), Sn^{2+}, is a cation with a charge of +2. The anion, fluoride, has a change of −1. Because 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 **SnF_2**.

h. The polyatomic ion ammonium, NH_4^+, has a +1 charge and the polyatomic ion sulfate, SO_4^{2-}, has a −2 charge. To balance the charge, we need 2 NH_4^+ cations. The correct formula is **$(NH_4)_2SO_4$**.

i. Silver forms only a +1 ion. The perchlorate ion, ClO_4^-, has a charge of −1. Because the charges are numerically the same, the ions combine in a one-to-one ratio. The correct formula is **$AgClO_4$**.

j. This is a molecular compound. The Greek prefixes tell you the number of each type of atom in the molecule: no prefix indicates 1 and tri- indicates 3. The correct formula is **BCl_3**.

2.64 a. CuCN

b. $Sr(ClO_2)_2$

c. $HBrO_4(aq)$

d. $HI(aq)$

e. $Na_2(NH_4)PO_4$

f. KH_2PO_4

g. IF_7

h. P_4S_{10}

i. HgO

j. Hg_2I_2

k. SeF_6

2.65 a. $Mg(NO_3)_2$ b. Al_2O_3 b. LiH b. Na_2S

2.66 a. one green sphere, one red sphere

b. one green sphere, two red spheres

c. three green spheres, two red spheres

d. two green spheres, one red sphere

2.67 A molecule is a combination of at least two atoms in a specific arrangement held together by electrostatic forces known as covalent chemical bonds.

2.68 An allotrope is one of two or more distinct forms of an element. For example, diamond and graphite are two allotropes of carbon. Allotropes have different chemical bonding of atoms of the same element. Isotopes have different nuclear structures.

2.69 A chemical formula denotes the composition of the substance.

a. 1:1

b. 1:3

c. 2:4 = 1:2

d. 4:6 = 2:3

2.70 A molecular formula shows the exact number of atoms of each element in a molecule. An empirical formula shows the lowest whole number ratio of the atoms of each element in a molecule.

2.71 Answers will vary. Example: C_2H_4 and C_4H_8

2.72 Organic compounds contain carbon and hydrogen, sometimes in combination with other elements such as oxygen, nitrogen, sulfur, and the halogens. Inorganic compounds generally do not contain carbon, although some carbon-containing species are considered inorganic.

2.73 Answers will vary.

Binary: carbon dioxide, CO_2

Ternary: dichloromethane, CH_2Cl_2

2.74 HCl in the gas phase is hydrogen chloride, a molecular compound. When dissolved in water, it dissociates completely into ions and is hydrochloric acid.

2.75 a. This is a polyatomic molecule that is an elemental form of the substance. It is not a compound.

b. This is a polyatomic molecule that is a compound.

c. This is a diatomic molecule that is a compound.

2.76 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.

2.77 **Elements:** N_2, S_8, H_2

Compounds: NH_3, NO, CO, CO_2, SO_2

2.78 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)

2.79 **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 a common 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 C_2N_2 is **CN**.

b. Dividing all subscripts by 6, the simplest whole number ratio of the atoms in C_6H_6 is **CH**.

c. The molecular formula as written, $\mathbf{C_9H_{20}}$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

d. Dividing all subscripts by 2, the simplest whole number ratio of the atoms in P_4O_{10} is $\mathbf{P_2O_5}$.

e. Dividing all subscripts by 2, the simplest whole number ratio of the atoms in B_2H_6 is $\mathbf{BH_3}$.

2.80 a. $AlBr_3$ b. $NaSO_2$ c. N_2O_5 d. $K_2Cr_2O_7$ e. HCO_2

2.81 $C_3H_7NO_2$

2.82 C_2H_6O (The formula for ethanol can also be written as C_2H_5OH or CH_3CH_2OH.)

2.83 a. nitrogen trichloride

b. iodine heptafluoride

c. tetraphosphorus hexoxide

d. disulfur dichloride

2.84 a. PBr_3 b. N_2F_4 c. XeO_4 d. SeO_3

2.85 All of these are molecular compounds. We use prefixes to express the number of each atom in the molecule. The molecular formulas and names are:

a. NF_3: nitrogen trifluoride

b. PBr_5: phosphorus pentabromide

c. SCl_2: sulfur dichloride

2.86 a. OF_2: oxygen difluoride

b. Al_2Br_6: dialuminum hexabromide

c. N_2F_4: dinitrogen tetrafluoride (also "perfluorohydrazine")

2.87 acid: compound that produces H^+; base: compound that produces OH^-; oxoacids: acids that contain oxygen; oxoanions: the anions that remain when oxoacids lose H^+ ions; hydrates: ionic solids that have water molecules in their formulas.

2.88 Uranium is radioactive. It loses mass because it constantly emits alpha (α) particles.

2.89 **(c)** 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.

2.90 The number of protons = 137 − 82 = 55. The element that contains 55 protons is cesium, Ce. There is one fewer electron than protons, so the charge of the cation is +1. The symbol for this cation is $\mathbf{Ce^{1+}}$.

2.91 Atomic number = 31 − 16 = 53. This anion has 15 protons, so it is a phosphorous ion. Since there are three more electron than protons, the ion has a −3 charge. The correct symbol is $\mathbf{P^{3-}}$.

2.92 a. Species with the same number of protons and electrons will be neutral. **A, F, G**.

b. Species with more electrons than protons will have a negative charge. **B, E**.

c. Species with more protons than electrons will have a positive charge. **C, D**.

d. **A**: ${}^{10}_{5}B$ **B**: ${}^{14}_{7}N^{3-}$ **C**: ${}^{39}_{19}K^{+}$ **D**: ${}^{66}_{30}Zn^{2+}$ **E**: ${}^{81}_{35}Br^{-}$ **F**: ${}^{11}_{5}B$ **G**: ${}^{19}_{9}F$

2.93 NaCl is an ionic compound; it doesn't consist of molecules.

2.94 **Yes**. The law of multiple proportions requires that the masses of sulfur combining with phosphorus must be in the ratios of small whole numbers. For the three compounds shown, four phosphorus atoms combine with three, seven, and ten sulfur atoms, respectively. If the atom ratios are in small whole number ratios, then the mass ratios must also be in small whole number ratios.

2.95 The species and their identification are as follows:

a. SO_2 molecule and compound

b. S_8 element and molecule

c. Cs element

d. N_2O_5 molecule and compound

e. O element

f. O_2 element and molecule

g. O_3 element and molecule

h. CH_4 molecule and compound

i. KBr compound, not molecule

j. S element

k. P_4 element and molecule

l. LiF compound, not molecule

2.96 a. This is an ionic compound. Prefixes are *not* used. The correct name is barium chloride.

b. Iron has a +3 charge in this compound. The correct name is iron(III) oxide.

c. NO_2^- is the nitrite ion. The correct name is cesium nitrite.

d. Magnesium is an alkaline earth metal, which always has a +2 charge in ionic compounds. The roman numeral is not necessary. The correct name is magnesium bicarbonate.

2.97 All masses are relative, which means that the mass of every object is compared to the mass of a standard object (such as the piece of metal in Paris called the "standard kilogram"). The mass of the standard object is determined by an international committee, and that mass is an arbitrary number to which everyone in the scientific community agrees.

Atoms are so small it is hard to compare their masses to the standard kilogram. Instead, we compare atomic masses to the mass of one specific atom. In the 19th century the atom was ^1H, and for a good part of the 20th century it was ^{16}O. Now it is ^{12}C, which is given the arbitrary mass of 12 amu exactly. All other isotopic masses (and therefore average atomic masses) are measured relative to the assigned mass of ^{12}C.

2.98 a. Ammonium is NH_4^+, not NH_3^+. The formula should be **$(NH_4)_2CO_3$**.

b. Calcium has a +2 charge and hydroxide has a –1 charge. The formula should be **$Ca(OH)_2$**.

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 **$ZnCr_2O_7$**.

2.99

Symbol	$^{11}_{5}B$	$^{54}_{26}Fe^{2+}$	$^{31}_{15}P^{3-}$	$^{196}_{79}Au$	$^{222}_{86}Rn$
Protons	5	26	15	79	86
Neutrons	6	28	16	117	136
Electrons	5	24	18	79	86
Net Charge	0	+2	–3	0	0

2.100 a. Ionic compounds are typically formed between metallic and nonmetallic elements.

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

2.101 a. Li^+, alkali metals always have a +1 charge in ionic compounds

b. S^{2-}

c. I^-, halogens have a −1 charge in ionic compounds

d. N^{3-}

e. Al^{3+}, aluminum always has a +3 charge in ionic compounds

f. Cs^+, alkali metals always have a +1 charge in ionic compounds

g. Mg^{2+}, alkaline earth metals always have a +2 charge in ionic compounds.

2.102 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?

2.103 The binary Group 7A element acids are: HF, hydrofluoric acid; HCl, hydrochloric acid; HBr, hydrobromic acid; HI, hydroiodic acid. Oxoacids containing Group 7A elements (using the specific examples for chlorine) are: $HClO_4$, perchloric acid; $HClO_3$, chloric acid; $HClO_2$, chlorous acid: HClO, hypochlorous acid.

Examples of oxoacids containing other Group A-block elements are: H_3BO_3, boric acid (Group 3A); H_2CO_3, carbonic acid (Group 4A); HNO_3, nitric acid and H_3PO_4, phosphoric acid (Group 5A); and H_2SO_4, sulfuric acid (Group 6A). Hydrosulfuric acid, H_2S, is an example of a binary Group 6A acid while HCN, hydrocyanic acid, contains both a Group 4A and 5A element.

2.104 a. C_2H_2, CH   a. C_6H_6, CH   a. C_2H_6, CH_3 a. C_3H_8, C_3H_8

2.105 a.

Isotope	$^{4}_{2}He$	$^{20}_{10}Ne$	$^{40}_{18}Ar$	$^{84}_{36}Kr$	$^{132}_{54}Xe$
No. Protons	2	10	18	36	54
No. Neutrons	2	10	22	48	78

b.

neutron/proton ratio	1.00	1.00	1.22	1.33	1.44

The neutron/proton ratio increases with increasing atomic number.

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

2.107 Cu, Ag, and Au are fairly chemically unreactive. This makes them especially suitable for making coins and jewelry that you want to last a very long time.

2.108 They generally do not react with other elements. Helium, neon, and argon are chemically inert.

2.109 Magnesium and strontium are also alkaline earth metals. You should expect the charge of the metal to be the same (+2). **MgO** and **SrO**.

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

2.111 a. $\dfrac{2\text{ red} : 1\text{ blue}}{1\text{ red} : 1\text{ blue}} = \mathbf{2:1}$

b. $\dfrac{1\text{ red} : 2\text{ blue}}{1\text{ red} : 1\text{ blue}} = \mathbf{1:2}$

c. $\dfrac{4\text{ red} : 2\text{ blue}}{1\text{ red} : 1\text{ blue}} = 4:2 = \mathbf{2:1}$

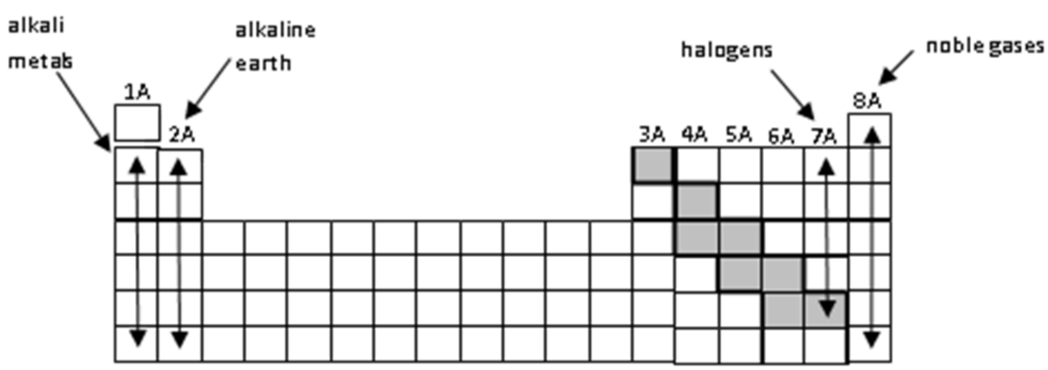
d. $\dfrac{5\text{ red} : 2\text{ blue}}{1\text{ red} : 1\text{ blue}} = \mathbf{5:2}$

2.113 The mass of fluorine reacting with hydrogen and deuterium would be the same. The ratio of F atoms to hydrogen (or deuterium) atoms is 1:1 in both compounds. This does not violate the law of definite proportions. When the law of definite proportions was formulated, scientists did not know of the existence of isotopes.

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

2.115 a. Br a. Rn a. Se a. Rb a. Pb

2.116



The metalloids are shown in gray.

2.117

Cation	Anion	Formula	Name
$\mathbf{Mg^{2+}}$	$\mathbf{HCO_3^-}$	$\mathbf{Mg(HCO_3)_2}$	Magnesium bicarbonate

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

2.118 a. $CO_2(s)$

b. NaCl

c. N_2O

d. $CaCO_3$

e. $NaHCO_3$

f. NH_3

g. H_2O

h. $Mg(OH)_2$

i. $MgSO_4 \cdot 7H_2O$

2.119 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

$$\Delta m = \frac{\Delta E}{c^2} = \frac{(1.715\times10^3\ \text{kJ})\left(\dfrac{1000\ \text{J}}{1\ \text{kJ}}\right)}{(2.998\times10^8\ \text{m/s})^2} = 1.908 \times 10^{-11}\ \text{kg} = \mathbf{1.908 \times 10^{-8}\ g}$$

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 1.908×10^{-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.

2.120 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:

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(3.04 \times 10^{-13} \text{ cm})^3 = 1.177 \times 10^{-37} \text{ cm}^3$$

The density of the nucleus can now be calculated.

$$\boldsymbol{d} = \frac{m}{V} = \frac{3.82 \times 10^{-23} \text{ g}}{1.177 \times 10^{-37} \text{ cm}^3} = \mathbf{3.25 \times 10^{14} \text{ g / cm}^3}$$

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:

$$11 \text{ electrons} \times \frac{9.1094 \times 10^{-28} \text{ g}}{1 \text{ electron}} = 1.00203 \times 10^{-26} \text{ g}$$

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:

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

$$V_{\text{atom}} = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(1.86 \times 10^{-8} \text{ cm})^3 = 2.695 \times 10^{-23} \text{ cm}^3$$

$$V_{\text{electrons}} = V_{\text{atom}} - V_{\text{nucleus}} = (2.695 \times 10^{-23} \text{ cm}^3) - (1.177 \times 10^{-37} \text{ cm}^3) = 2.695 \times 10^{-23} \text{ cm}^3$$

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

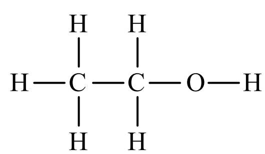
The density of the space occupied by the electrons can now be calculated.

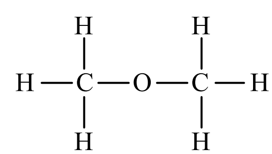
$$\boldsymbol{d} = \frac{m}{V} = \frac{1.00203 \times 10^{-26}\ \text{g}}{2.695 \times 10^{-23}\ \text{cm}^3} = \mathbf{3.72 \times 10^{-4}\ g / cm^3}$$

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.

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

2.122 Two different structural formulas for the molecular formula C_2H_6O are:





The second hypothesis of Dalton's Atomic Theory states that compounds are composed of atoms of more than one element, and in any given compound, the same types of atoms are always present in the same relative numbers. Both of the above compounds are consistent with the second hypothesis.

2.123 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

2.124 a. The following strategy can be used to convert from the volume of the Pt cube to the number of Pt atoms.

$$\text{cm}^3 \rightarrow \text{grams} \rightarrow \text{atoms}$$

$$1.0\text{ cm}^3 \times \frac{21.45\text{ g Pt}}{1\text{ cm}^3} \times \frac{1\text{ atom Pt}}{3.240 \times 10^{-22}\text{ g Pt}} = \mathbf{6.6 \times 10^{22}\ Pt\ atoms}$$

b. Since 74 percent of the available space is taken up by Pt atoms, 6.6×10^{22} atoms occupy the following volume:

$$0.74 \times 1.0\text{ cm}^3 = 0.74\text{ cm}^3$$

We are trying to calculate the radius of a single Pt atom, so we need the volume occupied by a single Pt atom.

$$\text{volume Pt atom} = \frac{0.74\text{ cm}^3}{6.6 \times 10^{22}\text{ Pt atoms}} = 1.12 \times 10^{-23}\text{ cm}^3/\text{Pt atom}$$

The volume of a sphere is $\frac{4}{3}\pi r^3$. Solving for the radius:

$$V = 1.12 \times 10^{-23}\text{ cm}^3 = \frac{4}{3}\pi r^3$$

$$r^3 = 2.67 \times 10^{-24}\text{ cm}^3$$

$$r = 1.4 \times 10^{-8}\text{ cm}$$

Converting to picometers:

$$\textbf{radius Pt atom} = (1.4 \times 10^{-8}\text{ cm}) \times \frac{0.01\text{ m}}{1\text{ cm}} \times \frac{1\text{ pm}}{1 \times 10^{-12}\text{ m}} = \mathbf{1.4 \times 10^2\ pm}$$

2.125 a. Assume that the nucleons (protons and neutrons) are hard objects of fixed size. Then the volume of the nucleus is well-approximated by the direct proportion $V = kA$, where A is the number of nucleons (mass number of the atom). For a spherical nucleus, then $V = kA = \frac{4}{3}\pi r^3$. Solving for r:

$$kA = \frac{4}{3}\pi r^3$$

$$\left(\frac{3}{4\pi}\right)kA = r^3$$

$$\left[\left(\frac{3}{4\pi}\right)kA\right]^{1/3} = r$$

$$\left[\left(\frac{3}{4\pi}\right)k\right]^{1/3}\left(A^{1/3}\right) = r$$

$$cA^{1/3} = r \quad (c \text{ is a constant})$$

b. For the volume calculation, use lithium-7 (A = 7).

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(r_0 A^{1/3}\right)^3 = \left(\frac{4}{3}\pi r_0^3\right)(A) = \left[\frac{4}{3}\pi\left(1.2\times10^{-15}\text{ m}\right)^3\right](7) \approx 5.1\times10^{-44}\text{ m}^3$$

c. Use r = 152 pm = 152 × 10^{-12} m for the atomic radius. Then, the atomic volume of lithium-7 is:

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(152\times10^{-12}\text{ m}\right)^3 \approx 1.5\times10^{-29}\text{ m}^3$$

The fraction of the atomic radius occupied by the nucleus is $\frac{5.1\times10^{-44}}{1.5\times10^{-29}} \approx 3.4\times10^{-15}$. This is consistent with Rutherford’s discovery that the nucleus occupies a very small region within the atom.

2.126 The average mass of the ball bearings is determined by the formula $D_a = \frac{m_a}{V_a}$ where D_a is the average density of the ball bearings, m_a is the average mass and V_a is the average volume. First, let's calculate the average volume of the bearings using the provided average diameter. The final answer is converted to units of cm^3 for subsequent density calculations.

$$V_a = \frac{4}{5}\pi r^3 = \frac{4}{5}\pi\left(\frac{4.175mm}{2}\right)^3 = 22.86mm^3 \times \left(\frac{1cm}{10mm}\right)^3 = 0.02286cm^3$$

Now let's determine the average density of the bearings. The density of copper is 8.92 g/cm^3 and the density of titanium is 4.507 g/cm^3. The percent alloy is calculated as $100\% - (30.4\% + 51.2\%) = 18.4\%$.

$$D_a = (8.92\, g/cm^3 \times 0.304) + (4.507\, g/cm^3 \times 0.512) + (5.659 g/cm^3 \times 0.184) = 6.061\, g/cm^3$$

Now the average mass can be calculated by rearranging the density equation such that $m_a = D_a \times V_a$.

$$m_a = D_a \times V_a = 6.061\, g/cm^3 \times 0.02286cm^3 = 0.1385g$$

Chapter Two: Atoms, Molecules, and Ions

Learning objectives

1. Identify all three hypotheses associated with Dalton's atomic theory and explain how they relate to the structure of matter and its associated reactions.
2. Recognize the importance of experiments conducted by Thomson, Millikan, Röntgen and Rutherford in regard to understanding the nature and structure of atoms.
3. Understand the different types of radiation that radioactive substances can produce.
4. Identify the location and physical properties of electrons, protons and neutrons in atoms.
5. Understand the nature and importance of isotopes.
6. Calculate the mass number of an isotope.
7. Utilize the mass number of an isotope to determine the number of electrons, protons or neutrons given other relevant information.
8. Utilize the periodic table to identify the chemical and physical properties of an element.
9. Categorize an element according to group based upon location in the periodic table.
10. Understand the nature of the atomic mass scale.
11. Calculate the average atomic mass of an element given the atomic mass and relative abundance of each of its naturally occurring isotopes.
12. Predict the charge of an ion formed from a main group element.
13. Name polyatomic ions and their associated charge.
14. Understand the information that chemical, molecular, and structural formulas provide.
15. Understand the differences between covalent and ionic bonding.
16. Determine the empirical formula of a compound given its molecular formula.
17. Utilize rules of nomenclature to name the different types of compounds including: covalent compounds, ionic compounds, oxoacids and hydrates.

Applications, Demonstrations, Tips and References

1. Page 38. Biological application.
2. Page 39. Instructor's Tip: If dimensional analysis did not identify the mathematically challenged students, calculating average atomic mass in this chapter will. Pointing out locations on campus where students can get help with their math skills may be helpful.
3. Page 40. Instructor's Tip: Historical footnotes can be used to bring meaning to some of the terms used in this chapter.
4. Page 40. Instructor's Tip: Nice lecture opener: Wouldn't it have been nice to study chemistry in ancient Greece? They only had four elements (earth, wind, water and fire).
5. Page 40. Literature: Mierrzecki, Roman."Dalton's atoms or Dalton's molecules? (SBS). *J. Chem. Educ.* **1981**, *58*, 1006.
6. Page 41. Organic application.
7. Page 42. Demonstration: You can easily demonstrate the law of conservation of mass using dry ice, an Erlenmeyer flask, scale and a balloon. The dry ice will sublime but no change in mass will occur.
8. Page 42. Multimedia: Law of conservation of mass.
9. Page 43. Instructor's Tip: Figures can be very helpful for illustrating the concept of an atom.

10. Page 43. Literature: Peake, Barrie M. "The discovery of the electron, proton, and neutron." *J. Chem. Educ.* **1989**, *66*, 738.
11. Page 43. Instructor's Tip: The word electron comes from the Greek word for amber. The term was introduced in 1891 by George Stoney.
12. Page 43. Multimedia: Cathode Ray Tube experiment
13. Page 44. Engineering application.
14. Page 44. Instructor's Tip: Consider talking about the production of various radioactive isotopes (iodine, cesium, etc.) that would be produced in a nuclear blast. This can be used to introduce the concept of isotopes and the fact that they just differ by the number of neutrons.
15. Page 45. Engineering application.
16. Page 45. Multimedia: Millikan Oil drop experiment
17. Page 46. Engineering application.
18. Page 47. Literature: Garrett, A. B. "The flash of genius 11. The neutron identified: Sir James Chadwick." *J. Chem. Educ.* **1962**, *39*, 638.
19. Page 48. Literature: Jensen, William B. "The Origins of the Symbols *A* and *Z* for Atomic Weight and Number." *J. Chem. Educ.* **2005**, *82*, 1764.
20. Page 48. Demonstration: Ellis, Arthur B. "Dramatizing isotopes: Deuterated ice cubes sink." *J. Chem. Educ.* **1990**, *67*, 159.
21. Page 48. Literature: Sein, Lawrence T., Jr. "Using Punnett Squares To Facilitate Students' Understanding of Isotopic Distributions in Mass Spectrometry." *J. Chem. Educ.* **2006**, *83*, 228.
22. Page 50. Literature: Laing, Michael. "The periodic table a new arrangement (PO)." *J. Chem. Educ.* **1989**, *66*, 746.
23. Page 50. Demonstration: Rizzo, Michelle M. et al. "Revisiting the Electric Pickle Demonstration." *J. Chem. Educ.* **2005**, *82*, 545.
24. Page 50. Literature: Hawkes, Stephen J. "Semimetallicity?" *J. Chem. Educ.* **2001**, *78*, 1686.
25. Page 50. Demonstration: Geselbracht, Margaret J. et al. "Mechanical Properties of Metals: Experiments with Steel, Copper, Tin, Zinc, and Soap Bubbles." *J. Chem. Educ.* **1994**, *71*, 254.
26. Page 50. Literature: Jensen, William B. "Why Helium Ends in "-ium."' *J. Chem. Educ.* **2004**, *81*, 944.
27. Page 51. Biological and environmental application.
28. Page 52. Engineering application.
29. Page 52. Literature: Last, Arthur M.; Webb, Michael J. "Using monetary analogies to teach average atomic mass (AA)." *J. Chem. Educ.* **1993**, *70*, 234.
30. Page 54. Literature: Schmid, Roland. "The Noble Gas Configuration --Not the Driving Force but the Rule of the Game in Chemistry." *J. Chem. Educ.* **2003**, *80*, 931.
31. Page 55. Instructor's Tip: It may help to stress why this system is not necessary for Group IA, IIA, IIIA metals (charge is predictable).
32. Page 56. Instructor's Tip: It is not necessary to make your students memorize all of the polyatomic ions, but you should clearly indicate which ones they will be responsible for.
33. Page 60. Literature: Schaeffer, Richard W et al. "Preparation and Analysis of Multiple Hydrates of Simple Salts." *J. Chem. Educ.* **2000**, *77*, 509.

34. Page 61. Literature: Jensen, William B. “The Origin of the Term Allotrope.” *J. Chem. Educ.* **2006**, *83*, 838.
35. Page 61. Organic application.
36. Page 62. Literature: Chimeno, Joseph. "How to Make Learning Chemical Nomenclature Fun, Exciting, and Palatable." *J. Chem. Educ.* **2000**, *77*, 144.
37. Page 62. Organic application.
38. Page 63. Organic application.
39. Page 64. Instructor’s Tip: You can ask students to name some common acids and bases to stimulate their interest.
40. Page 65. Literature: Meek, Terry L. "Acidities of oxoacids: Correlation with charge distribution." *J. Chem. Educ.* **1992**, *69*, 270.
41. Page 65. Instructor’s Tip: If students learn the rules for naming oxoacids, it makes it easier to remember some of the polyatomic ions (e.g. sulfite and sulfate).
42. Page 65. Organic application.
43. Page 65. Literature: Byrd, Shannon. “Learning the Functional Groups: Keys to Success.” *J. Chem. Educ.* **2001**, *78*, 1355.
44. Page 69. Instructor’s Tip: Constructing a nomenclature flowchart can help students learn how to name compounds.

End of Chapter Problems sorted by difficulty

Easy
1, 2, 3, 4, 7, 8, 9, 10, 11, 12, 13, 16, 17, 18, 19, 20, 21, 22, 23, 24, 25, 26, 27, 28, 29, 30, 31, 32, 33, 34, 35, 37, 38, 40, 41, 42, 43, 48, 49, 50, 51, 52, 53, 57, 58, 59, 60, 65, 66, 67, 68, 69, 73, 74, 75, 76, 77, 78, 79, 80, 81, 82, 87, 88, 93, 97, 100, 101, 102, 108, 109

Medium
5, 6, 14, 15, 39, 44, 45, 46, 47, 54, 55, 56, 61, 62, 63, 64, 70, 71, 72, 83, 84, 85, 86, 89, 90, 91, 92, 94, 95, 96, 98, 99, 103, 104, 105, 106, 107, 110, 111, 112, 114, 115, 116, 117, 119, 120, 121, 122, 123, 125, 127

Difficult
36, 113, 118, 124, 126, 128

End of Chapter Problems sorted by type

Review
1, 2, 3, 4, 6, 5, 7, 8, 9, 10, 11, 12, 13, 16, 17, 18, 19, 29, 30, 31, 32, 33, 40, 41, 42, 43, 50, 51, 52, 53, 54, 67, 68, 69, 70, 71, 72, 73, 74

Conceptual
4, 6, 20, 22, 24, 26, 55, 56, 57, 58, 59, 60, 61, 62, 63, 64, 65, 66, 75, 76, 77, 78, 79, 80, 81, 82, 83, 84, 85, 86, 88, 94, 113, 115, 116, 122

Biological
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