Chapter 2: Chemical Compounds

Teaching for Conceptual Understanding

The terms atom, molecule, element, and compound can be confusing to students and this confusion can persist beyond the introductory chemistry level. It is important to introduce each term clearly and to use many examples and non-examples. Ask students to draw nanoscale level diagrams of (1) atoms of an element, (2) molecules of an element, (3) atoms of a compound (NOT possible), and (4) molecules of a compound.

The idea of thinking about matter on different levels (macroscale, nanoscale, and symbolic) introduced in Chapter 1 can be reinforced by bringing to class a variety of element samples. When presenting each element, write the symbol, draw a nanoscale representation in its physical state at room temperature, and show a ball-and-stick or space-filled model.

Show students macroscopic samples of a variety of compounds together with a model, nanoscale level diagram, and symbol for each. Include both ionic and covalent compounds and be sure that all three states of matter are represented.

Students sometimes misinterpret the subscripts in chemical formulas. There are two common errors: (1) Assigning the subscript to the wrong element; for example, thinking that K_2S consists of 1 K atom and 2 S atoms instead of the reverse. (2) Not distributing quantities when parentheses are used; for example, thinking that $Mg(OH)_2$ has 1 O atom instead of 2 O atoms. It is important to test for these errors because they will lead to more mistakes when calculating molar masses, writing balanced chemical equations, and doing stoichiometric calculations.

Although we want our students to think conceptually and not rely on algorithms for problem solving, some algorithms, such as the one for naming compounds, are worth teaching to students. Two simple questions (and hints) students should ask themselves are: (1) Is there a metal in the formula? If not, prefixes will be used in the name. and (2) Does the metal form more than one cation? If yes, Roman numerals in parentheses will be used in the name.

A common misconception students have about ions, is that a positive ion has gained electrons and a negative ion has lost electrons. Use a tally of the protons, neutrons and electrons in an atom and the ion it forms to show students the basis of the charge of an ion.

The dissociation of an ionic compound into its respective ions when dissolved in water is another troublesome area. Research on student nanoscale diagrams has revealed misconceptions such as these: (1) $NiCl_2$ dissociates into Ni^{2+} and Cl_2, (2) NaOH dissociates into Na^+, O^{2-}, and H^+, (3) $Ni(OH)_2$ dissociates into Ni^{2+} and $(OH)_2{}^{2-}$. Having students draw nanoscale diagrams is an excellent way of testing their understanding of dissociation. Questions for Review and Thought 116, 117, 118 address this issue.

Some students completely ignore the charges when they look at formulas of ions; hence they see no difference between molecules and the ions with the same number and type of atoms, e.g., SO_3 and $SO_3{}^{2-}$ or NO_2 and $NO_2{}^-$. It is important to point out that an ion's formula is incomplete unless the proper charge is given, and that a substance is a compound if there is no charge specified.

Another means of assessing student understanding is to give them incorrect examples of a concept, term, or a problem's result and have them explain why it is incorrect. For some examples see Question for Review and Thought 128.

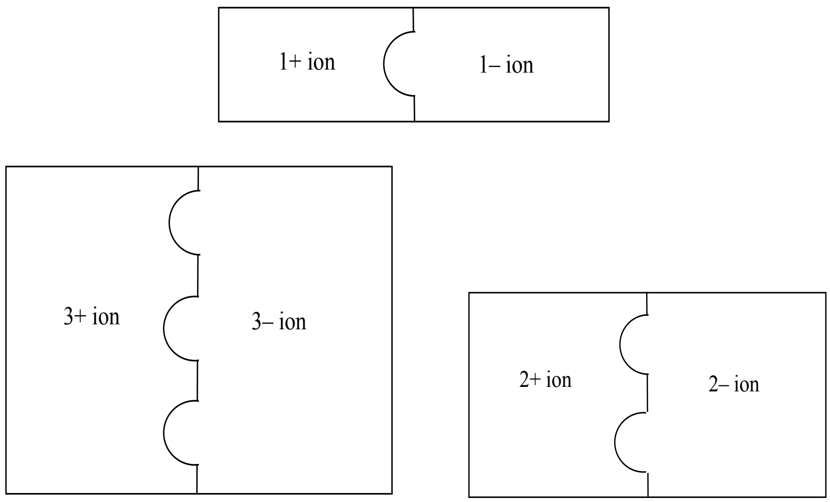
A quantity using the units of moles (Section 2-10) is one of the most important and most difficult concepts in a first course of introductory chemistry. Because Avogadro's number is so large, it is impossible to show students a 1-mole quantity of anything with distinct units that they can see and touch. Give numerous examples of mole quantities or have students think up some themselves, e.g., the Pacific Ocean holds one mole of teaspoons of water, or the entire state of Pennsylvania would be a foot deep in peas if one mole of peas were spread out over the entire state. Many students have the misconception that "mole" is a unit of mass; so, be on guard for this. Always ask students to identify the quantity, when you want them to calculate moles.

Suggestions for Effective Learning

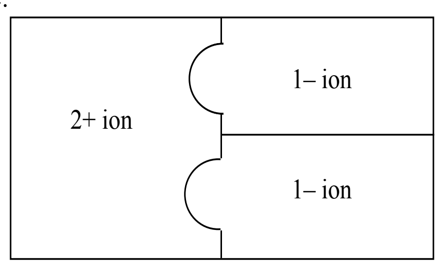
Many instructors skip organic and biochemistry topics in an introductory chemistry course because they think students will eventually take organic or biochemistry courses, so it is unnecessary to cover it in general chemistry.

The reality is that the majority of the students will not. Some of these students will be exposed to organic and biochemistry courses dealing with living systems (human, animal, or plant). The rest will leave college with a very limited view of chemistry. It is important to not skip the organic and biochemistry topics; they will add breadth to your course and spark student interest.

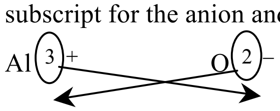
Not all college students are abstract thinkers; many are still at the concrete level when it comes to learning chemistry. The confusion surrounding the writing and understanding ionic chemical formulas can be eliminated by the use of simple jigsaw puzzle pieces. Below are templates for cation and anion cutouts. The physical manipulation or visualization of how these pieces fit together is enough for most students to grasp the concept.



Show them that the ions must fit together with all the notches paired. For example, magnesium chloride, is an example of a compound with a 2+ cation and two 1– anions.
The pieces fit together as shown here:



A tip for writing correct chemical formulas of ionic compounds is that the magnitude of charge on the cation is the subscript for the anion and vice versa. Consider the formula of the ionic compound made from Al^{3+} and O^{2-}:



results in the formula Al_2O_3.

Caution students that using this method can lead to incorrect formulas as in the case of Mg^{2+} and O^{2-} resulting in Mg_2O_2 instead of the correct formula of MgO.

Figure 2.6 shows the formation of an ionic compound.

In addition to showing representative samples of the compounds, it is useful to demonstrate examples of physical properties, e.g., cleaving a crystal with a sharp knife (Figure 2.9), conductivity of molten ionic compounds (Figure 2.10), etc.

When doing mole calculations, some students get answers that are off by a power of ten. While rare, this problem can usually be traced back to an error in a calculator entry. The student enters Avogadro's number as: 6.022, multiplication key, 10, exponent key, 23, which results in the value 6.022 x 10^{24}. It may be necessary to teach students how to correctly enter exponential numbers into their calculators.

In addition to showing representative samples of the elements, it is good to demonstrate physical properties, e.g., sublimation of iodine, to link back to material in Chapter 1 and chemical changes, e.g., Li, Na, and K reacting with water, to link to future material.

Cooperative Learning Activities

Questions, problems, and topics that can be used for Cooperative Learning Exercises and other group work are:

- Have students complete a matrix of names and/or formulas of compounds formed by specified cations and anions. This exercise can be used as a drill-and-practice or as an assessment of student knowledge prior to or after instruction. Use only those cations and anions most relevant to your course.

	Cl^-	O^{2-}	NO_3^-	PO_4^{3-}
Na^+	NaCl sodium chloride			
Fe^{2+}				
Fe^{3+}				
Al^{3+}				

Questions, problems, and topics that can be used for Cooperative Learning Exercises and other group work are:

- Have students list elements and compounds they interact with each day.
- Questions for Review and Thought from the end of this chapter: 4, 104-105, 114, 49-50, 126
- Conceptual Challenge Problems: CP2.A, CP2.B CP2.C, CP2.D, and CP2.E

End-of-Chapter Solutions for Chapter 2

Summary Problem

Part I

Result: **(a) Europium, Eu (b) Z = 63 (c) A = 151, (d) Atomic weight = 151.965 u (e) Lanthanide Series (f) metal (g) 1.78×10^9 m**

Analyze: Given the number of protons, and neutrons in the atom, determine the identity of the element, its symbol, the atomic number, the mass number, its location in the periodic table, and whether it is a metal, nonmetal or metalloid. Given the isotopic abundance and masses of two isotopes of this element, determine its atomic weight. Given the mass of the elements, determine the number of moles and the number of atoms. Given the diameter of atoms of this element, determine how many meters long a chain of this many atoms would be.

Plan and Execute:

(a) The atomic number is the same as the number of protons, 63. The number of protons is the same as the number of electrons in an uncharged atom. The element is Europium, with a symbol of Eu.

(b) Atomic number = Z = number of protons = 63.

(c) Isotope's mass number = A = Z + number of neutrons = 63 + 91 = 151.

(d) Calculate the weighted average of the isotope masses.

Every 10,000 atoms of the element contain 4,780 atoms of the ^{151}Eu isotope, with an atomic mass of 150.920 u, and 5,220 atoms of the ^{153}Eu isotope, with an atomic mass of 152.922 u.

$$\frac{4780.\text{atoms }^{151}\text{Eu}}{10000\ \text{Ag atoms}} \times \left(\frac{150.920\ \text{u}}{1\ \text{atom }^{151}\text{Eu}}\right) + \frac{5220.\ \text{atoms }^{153}\text{Eu}}{10000\ \text{Ag atoms}} \times \left(\frac{152.922\ \text{u}}{1\ \text{atom }^{153}\text{Eu}}\right) = 151.965\ \text{u/Eu atom}$$

(e) The element is found in the Lanthanide Series because it shows up in the row labeled "Lanthanides" (Atomic Number 58-71.

(f) This is a transition metal, according to the color-coding on the Periodic Table.

(e) Using the atomic weight calculated in (d), we can say that the molar mass is 151.965 g/mol.

$$1.25\ \text{mg Eu} \times \left(\frac{1\ \text{g}}{1000\ \text{mg}}\right) \times \left(\frac{1\ \text{mol Eu}}{151.965\ \text{g Eu}}\right) \times \left(\frac{6.022 \times 10^{23}\ \text{Eu atoms}}{1\ \text{mol Eu atoms}}\right) = 4.95 \times 10^{18}\ \text{Eu atoms}$$

The atomic radius is 180 pm. The diameter is twice the radius. Multiply the number of atoms by the atomic diameter, and use metric conversions to determine the number of meters.

$$4.95 \times 10^{18}\ \text{Eu atoms} \times 2 \times (180\ \text{pm}) \times \frac{1 \times 10^{-12}\ \text{m}}{1\ \text{pm}} = 1.78 \times 10^{9}\ \text{m}$$

☑ *Reasonable Result Check:* The periodic table lists an atomic weight very close to the same as the one calculated (151.964 u vs. 151.965 u). The number of moles is smaller than the number of grams. The number of atoms is very large. The length of the atom chain is quite long, considering the small size of each atom; however, given the large number of atoms it makes sense that the chain would be long.

Part II

Results: (a) **$(NH_4)_2Cr_2O_7$, NH_4NO_3**, and **$C_7H_5N_3O_6$** (b) **NH_4NO_3** (c) **$C_7H_5N_3O_6$** (d) none at room temperature; **$(NH_4)_2Cr_2O_7$ and NH_4NO_3** at elevated temperatures (e) **$(NH_4)_2Cr_2O_7$ and NH_4NO_3**

Analyze: Given information about three nitrogen-containing compounds, write formulas, compare mass percents of N, identify which compound has the lowest melting point, determine which conduct electricity at room temperature or at high temperature, and if they would conduct electricity in the molten state.

Plan and Execute:

(a) Ammonium dichromate contains the ammonium cation, NH_4^+ and the dichromate anion, $Cr_2O_7^{2-}$. The compound must have two +1 cations to balance the charge of the 2– anion, so its formula is $\mathbf{(NH_4)_2Cr_2O_7}$.

Ammonium nitrate contains the ammonium cation, NH_4^+ and the nitrate anion, NO_3^-. The compound must have one 1+ cation to balance the charge of the 1– anion, so its formula is $\mathbf{NH_4NO_3}$.

Trinitrotoluene contains seven carbon atoms, five hydrogen atoms, three nitrogen atoms, and six oxygen atoms, so its formula is $\mathbf{C_7H_5N_3O_6}$.

All three compounds are solids at room temperature and all three decompose at high temperatures.

(b) Calculate the mass of N in one mole of each compound, as well as the molar mass of each compound. Divide the mass of N by the molar mass of the compound and multiply by 100%.

One mole of $(NH_4)_2Cr_2O_7$ has two moles of N.

Mass of N in one mole of $(NH_4)_2Cr_2O_7$ = 2(14.0067 g/mol N) = 28.0134 g/mol N

Molar mass of $(NH_4)_2Cr_2O_7$ = 2(14.0067 g/mol N) + 8(1.0079 g/mol H)

+ 2(51.9961 g/mol Cr) + 7(15.9994 g/mol O) = 252.0646 g/mol

$$\%\ \mathrm{N} = \frac{\text{mass N per mol}}{\text{mass compound per mol}} \times 100\% = \frac{28.0134\ \text{g N}}{252.0646\ \text{compound}} \times 100\% = 11.1136\%\ \mathrm{N}$$

One mole of NH_4NO_3 has two moles of N.

Mass of N in one mole of NH_4NO_3 = 2(14.0067 g/mol N) = 28.0134 g N

Molar mass of NH_4NO_3 = 2(14.0067 g/mol N) + 4(1.0079 g/mol H) + 3(15.9994 g/mol O) = 80.0432 g/mol

$$\%\ \mathrm{N} = \frac{\text{mass N per mol}}{\text{mass compound per mol}} \times 100\% = \frac{28.0134\ \text{g N}}{80.0432\ \text{g compound}} \times 100\% = 34.9979\%\ \mathrm{N}$$

One mole of $C_7H_5N_3O_6$ has three moles of N.

Mass of N in one mole of $C_7H_5N_3O_6$ = 3(14.0067 g/mol N) = 42.0201 g/mol N

Molar mass of $C_7H_5N_3O_6$ = 7(12.0107 g/mol C) + 5(1.0079 g/mol H)

+ 3(14.0067 g/mol N) + 6(15.9994 g/mol O) = 227.1309 g/mol

$$\%\ \mathrm{N} = \frac{\text{mass N per mol}}{\text{mass compound per mol}} \times 100\% = \frac{42.0201\ \text{g N}}{227.1309\ \text{g compound}} \times 100\% = 18.5004\%\ \mathrm{N}$$

$\mathbf{NH_4NO_3}$ has the highest mass percent nitrogen.

(c) Molecular compounds have lower melting points than ionic compounds. The first two compounds, $(NH_4)_2Cr_2O_7$ and NH_4NO_3, are ionic, so we predict that $\mathbf{C_7H_5N_3O_6}$ has the lowest melting point.

(d) All three compounds are solid at room temperature, so we predict that **none** of them conduct electricity at room temperature. If temperature were high enough to melt the ionic compounds (without decomposing them), $\mathbf{(NH_4)_2Cr_2O_7}$ **and** $\mathbf{NH_4NO_3}$, the molten ions would conduct electricity. The molecular compound, $C_7H_5N_3O_6$, will not conduct electricity at any temperature.

(e) As described in (d), if the temperature of the ionic compounds, $\mathbf{(NH_4)_2Cr_2O_7}$ **and** $\mathbf{NH_4NO_3}$, could be raised high enough for melting, the molten ions would conduct electricity.

Part III

Results: **(a)** $\mathbf{Ca^{2+}}$**, 2+,** $\mathbf{PO_4^{3-}}$**, 3–, and** $\mathbf{F^-}$**, 1–;** $\mathbf{Na^+}$**, 1+ and** $\mathbf{S_2O_3^{2-}}$**, 2–;** $\mathbf{Ca^{2+}}$**, 2+ and** $\mathbf{C_2O_4^{2-}}$**, 2–**
(b) 504.303 g/mol of $\mathbf{Ca_5(PO_4)_3F}$**, 248.184 g/mol of** $\mathbf{Na_2S_2O_3 \cdot 5H_2O}$**, 164.127 g/mol of** $\mathbf{CaC_2O_4 \cdot 2H_2O}$

Analyze: Given the formulas of three compounds, identify the ions and their charges and calculate their molar masses.

Plan and Execute:

(a) Fluorapatite, $Ca_5(PO_4)_3F$, has calcium cations, Ca^{2+} with a charge of 2+, phosphate anion, PO_4^{3-} with a charge of 3–, and fluoride, F^- anion with a charge of 1–.

Hypo, $Na_2S_2O_3{\cdot}5H_2O$, has two sodium cations, Na^+ each with a charge of 1+ and a thiosulfate anion, $S_2O_3^{2-}$ with a charge of 2–.

Weddellite, $CaC_2O_4{\cdot}2H_2O$, has a calcium cation, Ca^{2+} with a charge of 2+ and oxalate anion, $C_2O_4^{2-}$ with a charge of 2–.

(b) Molar Mass of $Ca_5(PO_4)_3F$ = 5(40.078 g/mol Ca) + 3(30.9738 g/mol P)

+ 12(15.9994 g/mol O) + 18.9984 g/mol F = 504.303 g/mol of $Ca_5(PO_4)_3F$

Molar Mass of $Na_2S_2O_3{\cdot}5H_2O$ = 2(22.9898 g/mol Na) + 2(32.065 g/mol S)

+ 8(15.9994 g/mol O) + 10(1.0079 g/mol H) = 248.184 g/mol of $Na_2S_2O_3{\cdot}5H_2O$

Molar Mass of $CaC_2O_4{\cdot}2H_2O$ = (40.078 g/mol Ca) + 2(12.0107 g/mol C)

+ 6(15.9994 g/mol O) + 4(1.0079 g/mol H) = 164.127 g/mol of $CaC_2O_4{\cdot}2H_2O$

Part IV

Result: **(a) $C_6H_{13}O_3PS_2$ (b) $C_{12}H_{26}O_6P_2S_4$ (c) 1.259×10^{-4} mol (d) 7.583×10^{19} molecules**

Analyze: Given the mass percent composition of C H and O in a compound, the relationship between mass percent of S and P in the compound and the molar mass, determine the empirical formula, the molecular formula, the amount (moles) of compound in a sample with given mass, and the number of molecules in that sample.

Plan and Execute:

(a) The compound contains $C_xH_yO_zP_iS_j$, with 31.57% C, 5.74% H, and 21.03% O. The % S is 2.07 times % P.

Determine the percent that is not C, H, O, then use it to determine the actual %S and %P.

Set X = %S and Y = %P.

$$100.00\% = 31.57\%\ \text{C} + 5.74\%\ \text{H} + 21.03\%\ \text{O} + X + Y$$

$$X = 2.07Y$$

$$100.00\% - 31.57\%\ \text{C} - 5.74\%\ \text{H} - 21.03\%\ \text{O} = X + Y$$

$$41.66\% = X + Y = 2.07Y + Y = 3.07Y$$

$$Y = \frac{41.66\%}{3.07} = 13.6\%\ \text{P}$$

$$X = 2.07Y = 28.1\%\ \text{S}$$

A 100.00 g sample has 31.57 g C, 5.74 g H, 21.03 g O, 13.6 g P, and , 28.09 g S.

$$31.57\text{ g C} \times \left(\frac{1\text{ mol C}}{12.0107\text{ g C}}\right) = 2.628\text{ mol C} \qquad 5.74\text{ g H} \times \left(\frac{1\text{ mol H}}{1.0079\text{ g H}}\right) = 5.695\text{ mol H}$$

$$21.03\text{ g O} \times \left(\frac{1\text{ mol O}}{15.9994\text{ g O}}\right) = 1.314\text{ mol O} \qquad 13.6\text{ g P} \times \left(\frac{1\text{ mol P}}{30.9738\text{ g P}}\right) = 0.439\text{ mol P}$$

$$28.1\text{ g S} \times \left(\frac{1\text{ mol S}}{32.065\text{ g S}}\right) = 0.876\text{ mol S}$$

Set up mole ratio: 2.628 mol C : 5.695 mol H : 1.314 mol O : 0.439 mol P : 0.876 mol S

Simplify by dividing by 0.439 mol

$$6 \text{ mol C} : 13 \text{ mol H} : 3 \text{ mol O} : 1 \text{ P} : 2 \text{ S}$$

The empirical formula of dioxathion is $C_6H_{13}O_3PS_2$

(b) The molecular formula is $(C_6H_{13}O_3PS_2)_n$. Calculate the empirical formula molar mass, then divide it into the given molar mass for dioxathion (456.64 g/mol to determine n

Molar mass $C_6H_{13}O_3P_2S$ = 6(12.0107 g/mol C) + 13(1.0079 g/mol H)

+ 3(15.9994 g/mol O) + 30.9738 g/mol P + 2(32.065 g/mol S) = 228.269 g/mol

$$n = \frac{\text{molar mass comp}}{\text{molar mass emp. formula}} = \frac{456.64 \text{ g/mol}}{228.269 \text{ g/mol}} = 2$$

The molecular formula of dioxathion is $C_{12}H_{26}O_6P_2S_4$.

(c) Calculate the amount of dioxathion (in moles) in a sample using molar mass.

$$57.50 \text{ mg sample} \times \left(\frac{1 \text{ g}}{1000 \text{ mg}}\right) \times \left(\frac{1 \text{ mol}}{456.64 \text{ g}}\right) = 1.259 \times 10^{-4} \text{ mol}$$

(d) Use Avogadro's number to calculate the number of molecules of dioxathion in the sample.

$$1.259 \times 10^{-4} \text{ mol} \times \left(\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}}\right) = 7.583 \times 10^{19} \text{ molecules}$$

Questions for Review and Thought

Review Questions

1. *Result/Explanation:* The **coulomb** (C) is the fundamental unit of electrical charge.
2. *Result/Explanation:*
   (a) The proton is about **1800 times** heavier than the electron.
   (b) The charge on the proton has the **opposite sign** of the charge of the electron, but they have equal magnitude.
3. *Result/Explanation:* In a neutral atom, the number of protons is **equal** to the number of electrons.
4. *Result/Explanation:*
   (a) **Isotopes** of the same element have varying numbers of neutrons.
   (b) The mass number varies as the **number of neutrons** vary, since mass number is the sum of the number of protons and neutrons.
   (c) Answers to this question will vary. Common elements that have known isotopes are carbon (^{12}C and ^{13}C) and hydrogen (^{1}H, ^{2}H, and ^{3}H). Students may also give examples of boron, silicon, chlorine, magnesium, uranium, and neon based on the examples given in Section 2.3.
5. *Result/Explanation:*
   (a) (See Section 2.3) One **unified atomic mass unit**, symbol u (sometimes called amu), is exactly $\frac{1}{12}$ of the mass of one carbon-12 atom.
   (b) (End of Section 2.2) The **mass number** of an atom is the sum of the number of protons and the number of neutrons in the atom.
   (c) (See Section 2.11) The **molar mass** of any substance is the mass of one mole of that substance.
   (d) (See Section 2.3) Atoms of the same element that have different numbers of neutrons are called **isotopes**.

6. *Result/Explanation:* The "parts" that make up a chemical compound are atoms. Three pure (or nearly pure) compounds often encountered by people are: water (H_2O), table sugar (sucrose, $C_{12}H_{22}O_{11}$), and salt (NaCl).
A compound is different from a mixture because it has specific properties; the elements are present in definite proportions and can only be separated by chemical means. Mixtures can have variable properties and proportions, and the components of a mixture can be separated by physical means.

Topical Questions

Atomic Structure and Subatomic Particles (Section 2.1)

7. *Result/Explanation:* The masses and charges of the electron and proton are given in Section 2.1. Alpha particles are described as having two protons and two neutrons, so we double the charge of the proton to get the charge of the alpha particle. The alpha particle mass is not given in the textbook, but it is said to be the mass of one He^{2+} ion: mass of one He atom – 2(mass of electron)

$$\left(\frac{4.0026\text{ g}}{1\text{ mol}}\right)\times\left(\frac{1\text{ mole He atom}}{6.02214179\times10^{23}\text{ He atom}}\right)-2\times(9.1094\times10^{-28}\text{ g}) = 6.6447\times10^{-24}\text{ g/He}^{2+}$$

Name	Electric Charge (C)	Mass (g)	Deflected by Electric Field?
proton	1.6022×10^{-19}	1.6726×10^{-24}	**yes**
alpha particle	$\mathbf{3.2044\times10^{-19}}$	$\mathbf{6.6447\times10^{-24}}$	**yes**
electron	-1.6022×10^{-19}	$\mathbf{9.1094\times10^{-28}}$	**yes**

☑ *Reasonable Result Check:* The sum of two protons and two neutrons (6.6951×10^{-24} g) is slightly more than the mass of the alpha particle given at the National Institute of Standards and Technology web site: http://physics.nist.gov/cgi-bin/cuu/Value?mal .

8. *Result/Explanation:* The mass and charge of the neutron are given in Section 2.1.

Name	Electric Charge (C)	Mass (g)	Deflected by Electric Field?
neutron	$\mathbf{1.6022\times10^{-19}}$	1.6749×10^{-24}	**no**
gamma ray	**0**	0	**no**
beta ray	-1.6022×10^{-19}	$\mathbf{9.1094\times10^{-28}}$	**yes**

☑ *Reasonable Result Check:* See Section 2.1.

9. *Result:* **40,000 cm**

Analyze: If the nucleus is scaled to a diameter of a golf ball (4 cm), determine the diameter of the atom.

Plan: Find the accepted relationship between the size of the nucleus and the size of the atom. Use size relationships to get the diameter of the "artificially large" atom.

Execute:

From Figure 2.4, nucleus diameter is approximately 10^{-14} m and an atom's diameter is approximately 10^{-10} m

Determine atom-diameter/nucleus-diameter ratio: $\dfrac{10^{-10}\text{ m}}{10^{-14}\text{ m}} = 10^4$

So, the atom is about 10,000 times bigger than the nucleus. $10{,}000\times4\text{ cm} = 40{,}000\text{ cm}$

☑ *Reasonable Result Check:* A much larger nucleus means a much larger atom with a large atomic diameter.

10. *Result:* **the moon would be in the atom but the sun would not be**

Analyze, Plan, and Execute:

From http://oceanservice.noaa.gov/education/kits/tides/media/supp_tide02.html

The distance from the earth to the moon = 384,835 km. The distance from the earth to sun = 149,785,000 km

From http://www.universetoday.com/15055/diameter-of-earth/ , the average diameter of the Earth is 12,742 km.

From Figure 2.4, nucleus diameter is approximately 10^{-14} m and an atom's diameter is approximately 10^{-10} m

Determine atom-diameter/nucleus-diameter ratio: $\frac{10^{-10}\text{ m}}{10^{-14}\text{ m}} = 10^4$

Compare with moon-to-earth-distance/earth-diameter ratio:

$$\frac{384835\text{ km}}{12742\text{ km}} = 3.0 \times 10^1 < 10^4 \quad \text{Yes, the moon would be within the atom.}$$

Compare with sun-to-earth-distance/earth-diameter ratios:

$$\frac{149785000\text{ km}}{12742\text{ km}} = 1.2 \times 10^4 > 10^4 \quad \text{No, the sun would be outside the atom.}$$

☑ *Reasonable Result Check:* A much larger nucleus means a much larger atom with a large atomic diameter.

11. *Result:* (a) $^{67}_{34}Se$ (b) $^{72}_{36}Kr$ (c) $^{72}_{36}Kr$ (c) $^{67}_{34}Se$

Analyze and Plan: Given the symbol, $^{A}_{Z}S$, the number of neutrons is calculated with A – Z.

Execute:

For $^{67}_{34}Se$, the number of neutrons = 67 – 34 = 33. For $^{67}_{33}As$, the number of neutrons = 67 – 33 = 34.

For $^{67}_{35}Br$, the number of neutrons = 67 – 35 = 32. For $^{72}_{36}Kr$, the number of neutrons = 72 – 36 = 36.

(a) $^{67}_{34}Se$ contains 33 neutrons.

(b) $^{72}_{36}Kr$ contains the greatest number of neutrons

(c) $^{72}_{36}Kr$ contains equal number of protons and neutrons (36).

(d) Arsenic contains 33 protons. $^{67}_{34}Se$ contains 33 neutrons.

12. *Result:* (a) $^{115}_{50}Sn$ (b) $^{112}_{51}Sb$ (c) $^{115}_{49}In$ (c) $^{112}_{50}Sn$

Analyze and Plan: Given the symbol, $^{A}_{Z}S$, the number of neutrons is calculated A – Z.

Execute:

For $^{112}_{50}Sn$, the number of neutrons = 112 – 60 = 62. For $^{115}_{50}Sn$, the number of neutrons = 115 – 50 = 65.

For $^{112}_{51}Sb$, the number of neutrons = 112 – 51 = 61. For $^{115}_{49}In$, the number of neutrons = 115 – 49 = 66.

(a) $^{115}_{50}Sn$ contains 65 neutrons.

(b) $^{112}_{51}Sb$ contains the fewest number of neutrons

(c) $^{115}_{49}In$ contains the greatest number of neutrons.

(d) Sm atoms contain 62 protons. $^{112}_{50}Sn$ contains 62 neutrons.

Tools of Chemistry (Section 2.2 and 2.3)

13. *Result/Explanation:*

In Section 2.2, page 56, the **scanning tunneling microscope** (STM) is described. It has a metal probe in the shape of a needle with an extremely fine point that is brought extremely close to examine the sample surface. When the tip is close enough to the sample, electrons jump between the probe and the sample. The size and direction of this electron flow (the current) depend on the applied voltage, the distance between probe tip and sample, and the identity and location of the nearest sample atom on the surface and its closest neighboring atoms.

In Section 2.3, page 61, in a **mass spectrometer** atoms or molecules in a gaseous sample pass through a stream of high-energy electrons. Collisions between the electrons and the sample particles produce positive ions, mostly with 1+ charge.

14. *Result/Explanation:* In Section 2.3, page 61, a "Tools of Chemistry" box explains the Mass Spectrometer. The species that is moving through a mass spectrometer during its operation are **ions** (usually +1 cations) that have been formed from the sample molecules by a bombarding electron beam.

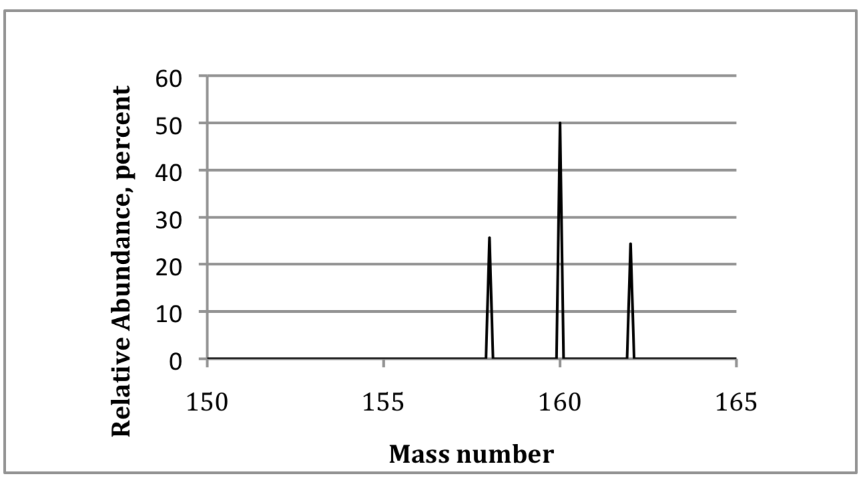
15. *Result/Explanation:* In Section 2.3, page 61, a "Tools of Chemistry" box explains the Mass Spectrometer. In a mass spectrum the x-axis is the **mass** of the ions and the y-axis is the **abundance** of the ions. The mass spectrum is a representation of the masses and abundances of the ions formed in the mass spectrometer, which are directly related to the molecular structure of the sample atoms or molecules.

16. *Result/Explanation:* The diatomic bromine molecule will be composed of: ^{79}Br–^{79}Br, ^{79}Br–^{81}Br, ^{81}Br–^{79}Br, and ^{81}Br–^{81}Br, causing peaks at mass numbers 158, 160, and 162.

Because the peak at 158 contains only bromine-79 and the relative abundance of that isotope is 50.69%, this peak has a relative abundance of (0.5069) × (0.5069)=0.2569, or 25.69%.

Because the peak at 162 contains only bromine-81 and the relative abundance of that isotope is 49.31%, this peak has a relative abundance of (0.4931) × (0.4931)=0.2569, or 24.31%.

Two isotopic combinations will have a peak at 160: ^{79}Br–^{81}Br and ^{81}Br–^{79}Br. This peak has a relative abundance of 2 × (0.5069) × (0.4931)=0.4999, or 49.99%

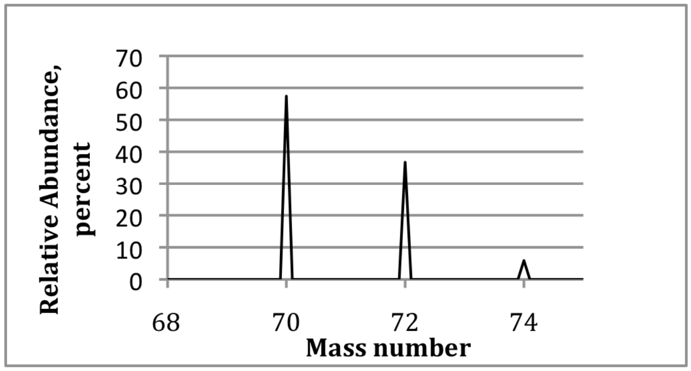


17. *Result/Explanation:* The diatomic chlorine molecule will be composed of: ^{35}Cl–^{35}Cl, ^{35}Cl–^{37}Cl, ^{37}Cl–^{35}Cl, and ^{37}Cl–^{37}Cl, causing peaks at mass numbers of 70, 72, and 74.

Because the peak at 70 contains only chlorine-35 and the relative abundance of that isotope is 75.77%, this peak has a relative abundance of (0.7577) × (0.7577)=0.5741, or 57.41%.

Because the peak at 74 contains only chlorine-37 and the relative abundance of that isotope is 24.23%, this peak has a relative abundance of (0.2423) × (0.2423)=0.0587, or 5.87%.

Two isotopic combinations will have a peak at 72: ^{35}Cl—^{37}Cl and ^{37}Cl—^{35}Cl. This peak has a relative abundance of 2 × (0.7577) × (0.2423)=0.3672, or 36.72%



Isotopes (Section 2-3)

18. *Result:* **number of neutrons, by three**

Analyze, Plan, and Execute: Uranium-235 differs from uranium-238 in terms of the number of neutrons in the atoms. Uranium-235 has three (238 – 235) fewer neutrons uranium-237.

19. *Result:* **number of neutrons, by two**

Analyze, Plan, and Execute: Strontium-90 differs from strontium-88 in terms of the number of neutrons in the atoms. Strontium-90 has two (90– 88) fewer neutrons strontium-88.

20. *Result:* **27 electrons, 27 protons, and 33 neutrons**

Analyze, Plan, and Execute: Given the identity of an element (cobalt) and the atom's mass number (60), find the number of electrons, protons, and neutrons in the atom.

Look up the symbol for cobalt and find that symbol on the periodic table. The periodic table gives the atomic number. The atomic number is the number of protons. The number of electrons is equal to the number of protons since the atom has no charge. The number of neutrons is the difference between the mass number and the atomic number.

The element cobalt has the symbol Co. On the periodic table, we find it listed with the atomic number 27. So, the atom has 27 protons, 27 electrons and (60 – 27 =) 33 neutrons.

☑ *Reasonable Result Check:* The number protons and electrons must be the same (27=27). The sum of the protons and neutrons is the mass number (27 + 33 = 60).

21. *Result:* **43 protons**, **43 electrons and 56 neutrons**

Analyze: Given the identity of an element (technetium) and the atom's mass number (99), find the number of electrons, protons, and neutrons in the atom.

Plan: Look up the symbol for technetium and find that symbol on the periodic table. (Refer to the strategy described in Question 20 for details.)

Execute: The element technetium has the symbol Tc. On the periodic table, we find it listed with the atomic number 43. So, the atom has 43 protons, 43 electrons and (99 – 43 =) 56 neutrons.

☑ *Reasonable Result Check:* The number protons and electrons must be the same (43=43). The sum of the protons and neutrons is the mass number (56 + 43 = 99).

22. *Result:* **78.92 u/atom**

Analyze: Given the average atomic weight of an element and the percentage abundance of one isotope, determine the atomic weight of the only other isotope.

Plan: Using the fact that the sum of the percents must be 100%, determine the percent abundance of the second isotope. Knowing that the weighted average of the isotope masses must be equal to the reported atomic weight, set up a relationship between the known atomic mass and the various isotope masses using a variable to describe the second isotope's atomic weight.

Execute: We are told that the natural abundance of ^{81}Br is 49.31% and that there are only two isotopes. To calculate the percent abundance of the other isotope, subtract from 100%:

$$100.0\% - 49.31\% = 50.69\%$$

These percentages tell us that every 10000 atoms of bromine contain 4931 atoms of the ^{81}Br isotope and 5069 atoms of the other bromine isotope *(limited to 4 sig figs)*. The atomic weight for Br is given as 79.904 u/atom. The isotopic mass of ^{81}Br isotope is 80.916289 u/atom. Let X be the atomic mass of the other isotope of bromine.

$$\frac{4931 \text{ atoms } ^{81}\text{Br}}{10000 \text{ Br atoms}} \times \left(\frac{80.916289 \text{ amu}}{1 \text{ atom } ^{81}\text{Br}}\right) + \frac{5069 \text{ atoms other isotope}}{10000 \text{ Br atoms}} \times \left(\text{X} \frac{\text{amu}}{\text{atom}}\right) = 79.904 \frac{\text{amu}}{\text{Br atom}}$$

Solve for X

$$39.90 + 0.5069 \text{ X} = 79.904$$

$$\text{X} = 78.92 \text{ u/atom} \quad \textit{(limited to 4 sig figs)}$$

☑ *Reasonable Result Check:* Because the relative abundance is very close to 50% for each isotope, we expected the mass of the lighter isotope to be lower than the mass of the heavier isotope.

23. *Result:* **11.01 u/atom**

Analyze: Given the average atomic weight of an element and the percentage abundance of one isotope, determine the atomic weight of the only other isotope.

Plan: Using the fact that the sum of the percents must be 100%, determine the percent abundance of the second isotope. Knowing that the weighted average of the isotope masses must be equal to the reported atomic weight, set up a relationship between the known atomic mass and the various isotope masses using a variable to describe the second isotope's atomic weight.

Execute: We are told that natural boron is 19.91% ^{10}B and that there are only two isotopes. To calculate the percent abundance of the other isotope, subtract from 100%:

$$100.00\% - 19.91\% = 80.09\%$$

These percentages tell us that every 10000 atoms of boron contain 1991 atoms of the ^{10}B isotope and 8009 atoms of the other boron isotope *(limited to 4 sig figs)*. The atomic weight for B is given as 10.811 u/atom. In Section 2-3a, the isotopic mass of ^{10}B isotope is given as 10.0129 u/atom. Let X be the atomic mass of the other isotope of boron.

$$\frac{1991 \text{ atoms } ^{10}\text{B}}{10000 \text{ B atoms}} \times \left(\frac{10.0129 \text{ u}}{1 \text{ atom } ^{10}\text{B}}\right) + \frac{8009 \text{ atoms other isotope}}{10000 \text{ B atoms}} \times \left(\text{X} \frac{\text{u}}{\text{atom}}\right) = 10.811 \frac{\text{u}}{\text{B atom}}$$

Solve for X

$$1.994 + 0.8009 \text{ X} = 10.811$$

$$\text{X} = 11.01 \text{ u/atom} \quad \textit{(limited to 4 sig figs)}$$

☑ *Reasonable Result Check:* Section 2-3a gives the atomic weight of ^{11}B to be 11.0093, which is the same as the answer here, within the given significant figures.

24. *Result:* **(a)** $^{23}_{11}\mathbf{Na}$ **(b)** $^{39}_{18}\mathbf{Ar}$ **(c)** $^{69}_{31}\mathbf{Ga}$

Analyze: Given the identity of an element and the number of neutrons in the atom, determine the atomic symbol $^{A}_{Z}X$.

Plan: Look up the symbol for the element and find that symbol on the periodic table. The periodic table gives the atomic number (Z), which represents the number of protons. Add the number of neutrons to the number of protons to get the mass number (A).

Execute:

(a) The element sodium has the symbol Na. On the periodic table, we find it listed with the atomic number 11. The given number of neutrons is 12. So, (11 + 12 =) 23 is the mass number for this sodium atom. Its atomic symbol looks like this: $^{23}_{11}\mathbf{Na}$.

(b) The element argon has the symbol Ar. On the periodic table, we find it listed with the atomic number 18. The given number of neutrons is 21. So, (18 + 21 =) 39 is the mass number for this argon atom. Its atomic symbol looks like this: $\mathbf{^{39}_{18}Ar}$.

(c) The element gallium has the symbol Ga. On the periodic table, we find it listed with the atomic number 31. The given number of neutrons is 38. So, (31 + 38 =) 69 is the mass number for this gallium atom. Its atomic symbol looks like this: $\mathbf{^{69}_{31}Ga}$.

☑ *Reasonable Result Check:* Mass number should be close to (but not exactly the same as) the atomic weight also given on the periodic table. Sodium's atomic weight (22.99) is close to the 23 mass number. Argon's atomic weight (39.95) is close to the 39 mass number. Gallium's atomic weight (69.72) is close to the 69 mass number.

25. *Result:* **(a)** $\mathbf{^{15}_{7}N}$ **(b)** $\mathbf{^{64}_{30}Zn}$ **(c)** $\mathbf{^{129}_{54}Xe}$

Analyze: Given the identity of an element and the number of neutrons in the atom, determine the atomic symbol $^{A}_{Z}X$.

Plan: Look up the symbol for the element and find that symbol (X) on the periodic table. The periodic table gives the atomic number (Z), which represents the number of protons. Add the number of neutrons to the number of protons to get the mass number (A).

Execute:

(a) The element nitrogen has the symbol N. On the periodic table, we find it listed with the atomic number 7. The given number of neutrons is 8. So, (7 + 8 =) 15 is the mass number for this nitrogen atom. Its atomic symbol looks like this: $\mathbf{^{15}_{7}N}$.

(b) The element zinc has the symbol Zn. On the periodic table, we find it listed with the atomic number 30. The given number of neutrons is 34. So, (30 + 34 =) 64 is the mass number for this zinc atom. Its atomic symbol looks like this: $\mathbf{^{64}_{30}Zn}$.

(c) The element xenon has the symbol Xe. On the periodic table, we find it listed with the atomic number 54. The given number of neutrons is 75. So, (54 + 75 =) 129 is the mass number for this xenon atom. Its atomic symbol looks like this: $\mathbf{^{129}_{54}Xe}$.

☑ *Reasonable Result Check:* Mass number should be close to (but not exactly the same as) the atomic weight also given on the periodic table. Nitrogen's atomic weight (14.01) is close to the 15 mass number. Zinc's atomic weight (65.38) is close to the 64 mass number. Xenon's atomic weight (131.3) is close to the 129 mass number.

26. *Result:* **See calculation below**

Analyze: Using the exact mass and the percent abundance of several isotopes of an element, determine the atomic weight.

Plan: Calculate the weighted average of the isotope masses.

Execute: Every 100 atoms of lithium contains 7.500 atoms of the ^{6}Li isotope and 92.50 atoms of the ^{7}Li isotope.

$$\frac{7.500 \text{ atoms } ^{6}\text{Li}}{100 \text{ Li atoms}} \times \left(\frac{6.015121 \text{ u}}{1 \text{ atom } ^{6}\text{Li}}\right) + \frac{92.50 \text{ atoms } ^{7}\text{Li}}{100 \text{ Li atoms}} \times \left(\frac{7.016003 \text{ u}}{1 \text{ atom } ^{7}\text{Li}}\right) = 6.941 \text{ u/Li atom}$$

☑ *Reasonable Result Check:* The periodic table value for atomic weight is the same as calculated here.

27. *Result:* **See calculation below**

Analyze: Using the exact mass and the percent abundance of several isotopes of an element, determine the atomic weight.

Plan: Calculate the weighted average of the isotope masses.

Execute: Every 10000 atoms of magnesium contains 7899 atoms of the ^{24}Mg isotope, 1000 atoms of the ^{25}Mg isotope, and 1101 atoms of the ^{26}Mg isotope.

$$\frac{7899 \text{ atoms } ^{24}\text{Mg}}{10000 \text{ Mg atoms}} \times \left(\frac{23.985042 \text{ u}}{1 \text{ atom } ^{24}\text{Mg}}\right) + \frac{1000 \text{ atoms } ^{25}\text{Mg}}{10000 \text{ Mg atoms}} \times \left(\frac{24.98537 \text{ u}}{1 \text{ atom } ^{25}\text{Mg}}\right)$$

$$+ \frac{1101 \text{ atoms } ^{26}\text{Mg}}{10000 \text{ Mg atoms}} \times \left(\frac{25.982593 \text{ u}}{1 \text{ atom } ^{26}\text{Mg}}\right) = 24.31 \frac{\text{u}}{\text{Mg atom}}$$

Notice: The given percentages limit each term to four significant figures, therefore the first term has only two decimal places. So, this answer is rounded off significantly.

☑ *Reasonable Result Check:* The periodic table value for atomic weight is the same as calculated here, within the limits of uncertainty

28. *Result:* **60.12% ^{69}Ga, 39.88% ^{71}Ga**

Analyze: Using the exact mass of two isotopes and the atomic weight, determine the abundance of the isotopes.

Plan: Establish variables describing the isotope percentages. Set up two relationships between these variables. The sum of the percents must be 100%, and the weighted average of the isotope masses must be the reported atomic mass.

Execute: Set X% ^{69}Ga and Y% ^{71}Ga. This means: Every 100 atoms of gallium contain X atoms of the ^{69}Ga isotope and Y atoms of the ^{71}Ga isotope.

$$\frac{\text{X atoms } ^{69}\text{Ga}}{100 \text{ Ga atoms}} \times \left(\frac{68.9257 \text{ u}}{1 \text{ atom } ^{69}\text{Ga}}\right) + \frac{\text{Y atoms } ^{71}\text{Ga}}{100 \text{ Ga atoms}} \times \left(\frac{70.9249 \text{ u}}{1 \text{ atom } ^{71}\text{Ga}}\right) = 69.723 \frac{\text{u}}{\text{Ga atom}}$$

And, X + Y = 100%. We now have two equations and two unknowns, so we can solve for X and Y algebraically. Solve the first equation for Y: Y = 100 – X. Plug that in for Y in the second equation. Then solve for X:

$$\frac{X}{100} \times (68.9257) + \frac{100 - X}{100} \times (70.9249) = 69.723$$

$$0.689257X + 70.9249 - 0.709249X = 69.723$$

$$70.9249 - 69.723 = 0.709249X - 0.689257X = (0.709249 - 0.689257)X$$

$$1.202 = (0.019992)X$$

X = 60.12, so there is 60.12% ^{69}Ga

Now, plug the value of X in the first equation to get Y.

Y = 100 – X = 100 – 60.12 = 39.88, so there is 39.88% ^{69}Ga

Therefore the abundances for these isotopes are: 60.12% ^{69}Ga and 39.88% ^{71}Ga

☑ *Reasonable Result Check:* The periodic table value for the atomic weight is closer to 68.9257 than it is to 70.9249, so it makes sense that the percentage of ^{69}Ga is larger than ^{71}Ga. The sum of the two percentages is 100.00%.

29. *Result:* **39.95 u/atom**

Analyze: Knowing that almost all of the argon in nature is ^{40}Ar, a good estimate for the atomic weight of argon is a little less than 40 u/atom. Using the exact mass and the percent abundance of several isotopes of an element, determine the atomic weight.

Plan: Calculate the weighted average of the isotope masses.

Execute: Every 100000 atoms of argon contains 337 atoms of the ^{36}Ar isotope, 63 atoms of the ^{38}Ar isotope, and 99600 atoms of the ^{40}Ar isotope.

$$\frac{337 \text{ atoms } ^{36}\text{Ar}}{100000 \text{ Ar atoms}} \times \left(\frac{35.968 \text{ u}}{1 \text{ atom } ^{36}\text{Ar}}\right) + \frac{63 \text{ atoms } ^{38}\text{Ar}}{100000 \text{ Ar atoms}} \times \left(\frac{37.963 \text{ u}}{1 \text{ atom } ^{38}\text{Ar}}\right)$$

$$+ \frac{99600 \text{ atoms } ^{40}\text{Ar}}{100000 \text{ Ar atoms}} \times \left(\frac{39.962 \text{ u}}{1 \text{ atom } ^{40}\text{Ar}}\right) = 39.95 \frac{\text{u}}{\text{Ar atom}}$$

☑ *Reasonable Result Check:* This calculated answer matches the estimate. Also, the periodic table value for the atomic weight is the same as this calculated value.

Ions and Ionic Compounds (Section 2-4)

30. *Result:* **(a) Li^+ (b) Sr^{2+} (c) Al^{3+} (d) Zn^{2+}**

Analyze and Plan: A general rule for the charge on a metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation's positive charge.

Execute:

(a) Lithium (Group 1A) Li^+

(b) Strontium (Group 2A) Sr^{2+}

(c) Aluminum (Group 3A) Al^{3+}

(d) Zinc (Group 2B) Zn^{2+}

31. *Result:* **(a) N^{3-} (b) S^{2-} (c) Cl^- (d) I^-**

Analyze and Plan: For nonmetal elements in Groups 5A-7A, the electrons gained by an atom to form a stable anion are calculated using the formula: 8 – (group number). That means the (group number) – 8 is the negative charge of the anion.

Execute:

(a) Nitrogen (Group 5A) 5 – 8 = –3 N^{3-}

(b) Sulfur (Group 6A) 6 – 8 = –2 S^{2-}

(c) Chlorine (Group 7A) 7 – 8 = –1 Cl^-

(d) Iodine (Group 7A) 7 – 8 = –1 I^-

32. *Result:* **(a) 2+ (b) 3– (c) 2+ or 3+ (d) 2–**

Analyze and Plan: A general rule for the charge on a monatomic metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation's positive charge. Transition metals often have a +2 charge. Some have +3 and +1 charged ions, as well. For nonmetal elements in Groups 5A-7A, the electrons gained by an atom to form a stable monatomic anion are calculated by subtracting the group number from 8. The difference between the group number and 8 is the negative charge of the anion.

Execute:

(a) Magnesium (Group 2A) has a 2+ charge. Mg^{2+}

(b) Phosphorus (Group 5A) 5 – 8 = –3 P^{3-}

(c) Iron (a transition metal) has a 2+ or 3+ charge. Fe^{2+} or Fe^{3+}

(d) Selenium (Group 6A) 6 – 8 = –2 Se^{2-}

33. *Result:* **(a) 3+ (b) 1– (c) 1+ (d) 3–**

Analyze and Plan: A general rule for the charge on a monatomic metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation's positive charge. Transition metals often have a +2 charge. Some have +3 and +1 charged ions, as well. For nonmetal elements in Groups 5A-7A, the electrons gained by an atom to form a stable monatomic anion are calculated using the formula: 8 – (group number). That means the (group number) – 8 is the negative charge of the anion.

Execute:

(a) Gallium (Group 3A) has a 3+ charge. Ga^{3+}

(b) Fluorine (Group 7A) 7 – 8 = –1 F^{-}

(c) Silver (a transition metal, Group 1B) has a 1+ charge. Ag^{+}

(d) Nitrogen (Group 5A) 5 – 8 = –3 N^{3-}

34. *Result:* **CoO, Co_2O_3**

Analyze, Plan, and Execute: Cobalt ions are Co^{2+} and Co^{3+}. Oxide ion is O^{2-}. The two compounds containing cobalt and oxide are made from the neutral combination of the charged ions:

One Co^{2+} and one O^{2-} [net charge = +2 + (–2) = 0] CoO

Two Co^{3+} and three O^{2-} [net charge = 2(+3) + 3(–2) = 0] Co_2O_3

35. *Result:* **$PbCl_2$, $PbCl_4$**

Analyze, Plan, and Execute: Two compounds containing lead and chloride are made from the neutral combination of the charged ions:

One Pb^{2+} and two Cl^{-} [net charge = +2 + 2(–1) = 0] $PbCl_2$

One Pb^{4+} and four Cl^{-} [net charge = +4 + 4(–1) = 0] $PbCl_4$

36. *Result:* **(c) and (d) are correct formulas. (a) $AlCl_3$, (b) NaF**

Analyze, Plan, and Execute:

(a) Aluminum ion (from Group 3A) is Al^{3+}. Chloride ion (from Group 7A) is Cl^{-}.

AlCl is **not** a neutral combination of these two ions. The correct formula would be $AlCl_3$.

[net charge = +3 + 3(–1) = 0]

(b) Sodium ion (Group 1A) is Na^{+}. Fluoride ion (from Group 7A) is F^{-}.

NaF_2 is **not** a neutral combination of these two ions. The correct formula would be NaF.

[net charge = +1 + (–1) = 0]

(c) Gallium ion (from Group 3A) is Ga^{3+}. Oxide ion (from Group 6A) is O^{2-}.

Ga_2O_3 **is** the correct neutral combination of these two ions.

[net charge = 2(+3) + 3(–2) = 0]

(d) Magnesium ion (from Group 2A) is Mg^{2+}. Sulfide ion (from Group 6A) is S^{2-}.

MgS **is** the correct neutral combination of these two ions.

[net charge = +2 + (–2) = 0]

37. *Result:* **(b) and (d) are correct formulas. (a) CaO, (c) FeO or Fe_2O_3**

Analyze, Plan, and Execute:

(a) Calcium ion (from Group 2A) is Ca^{2+}. Oxide ion (from Group 6A) is O^{2-}.

Ca_2O is **not** a neutral combination of these two ions. The correct formula would be CaO.

[net charge = +2 + (–2) = 0]

(b) Strontium ion (Group 2A) is Sr^{2+}. Chloride ion (from Group 7A) is Cl^{-}.

$SrCl_2$ **is** the correct neutral combination of these two ions. [net charge = +2 + 2(–1) = 0]

(c) Iron ion (from the transition elements) is Fe^{3+} or Fe^{2+}. Oxide ion (from Group 6A) is O^{2-}. Fe_2O_5 is **not** a neutral combination of these ions. The correct possible formulas would be

FeO [net charge = +2 + (–2) = 0] or Fe_2O_3 [net charge = 2(+3) + 3(–2) = 0]

(d) Potassium ion (from Group 1A) is K^+. Oxide ion (from Group 6A) is O^{2-}.

K_2O **is** the correct neutral combination of these two ions. [net charge = 2(+1) + (–2) = 0]

38. *Result:* **(b), (c), and (e) are ionic, because the compounds contain metals and nonmetals together**

Analyze and Plan: To tell if a compound is ionic or not, look for metals and nonmetals together, or common cations and anions. If a compound contains only nonmetals or metalloids and nonmetals, it is likely not ionic.

Execute:

(a) CF_4 contains only nonmetals. Not ionic.

(b) $SrBr_2$ has a metal and nonmetal together. Ionic.

(c) $Co(NO_3)_3$ has a metal and nonmetals together. Ionic.

(d) SiO_2 contains a metalloid and a nonmetal. Not ionic.

(e) KCN has a metal and nonmetals together. Ionic.

(f) SCl_2 contains only nonmetals. Not ionic.

39. *Result:* **Only (a) is ionic, with metal and nonmetal combined. (b)-(e) are composed of only non-metals.**

Analyze and Plan: To tell if a compound is ionic or not, look for metals and nonmetals together, or common cations and anions. If a compound contains only nonmetals or metalloids and nonmetals, it is likely not ionic.

Execute:

(a) NaH has a metal and a nonmetal together. Ionic.

(b) HCl contains only nonmetals. Not ionic.

(c) NH_3 contains only nonmetals. Not ionic.

(d) CH_4 contains only nonmetals. Not ionic.

(e) HI contains only nonmetals. Not ionic.

Naming Ions and Ionic Compounds (Section 2-5)

40. *Result:* **$BaSO_4$, barium ion, 2+, sulfate, 2–; $Mg(NO_3)_2$, magnesium ion, 2+, nitrate, 1–; $NaCH_3COO$, sodium ion, 1+, acetate, 1–**

Analyze, Plan, and Execute: Barium sulfate is $BaSO_4$. It contains a barium ion (Ba^{2+}), with a 2+ electrical charge, and a sulfate ion (SO_4^{2-}), with a 2– electrical charge. Magnesium nitrate is $Mg(NO_3)_2$. It contains a magnesium ion (Mg^{2+}), with a 2+ electrical charge, and two nitrate ions (NO_3^-), each with a 1– electrical charge. Sodium acetate is $NaCH_3COO$. It contains a sodium ion (Na^+), with a 1+ electrical charge, and an acetate ion (CH_3COO^-), with a 1– electrical charge. (Notice: Occasionally the Na^+ is written on the other end of the acetate formula like this CH_3COONa. It is done that way because the negative charge on acetate is on one of the oxygen atoms, so that's where the Na^+ cation will be attracted.)

41. *Result:* **$Ca(NO_3)_2$, calcium ion, 2+, nitrate , 1–; $BaCl_2$, barium ion, 2+, chloride, 1–; $(NH_4)_3(PO_4)$, ammonium ion, 1+, phosphate, 3–**

Analyze, Plan, and Execute: Calcium nitrate is $Ca(NO_3)_2$, barium chloride is $BaCl_2$, and ammonium phosphate is $(NH_4)_3(PO_4)$. The ions in calcium nitrate are Ca^{2+}, called calcium ion with a 2+ charge, and NO_3^-, called nitrate ion with a 1– charge. The ions in barium chloride are Ba^{2+}, called barium ion with a 2+ charge, and Cl^-, called chloride ion with a 1– charge. The ions in ammonium phosphate are NH_4^+, called ammonium ion with a

1+ charge, and PO_4^{3-}, called phosphate ion with a 3– charge.

42. *Result:* **(a) $Ni(NO_3)_2$ (b) $NaHCO_3$ (c) LiClO (d) $Mg(ClO_3)_2$ (e) $CaSO_3$**

Analyze, Plan, and Execute:

(a) Nickel(II) ion is Ni^{2+}. Nitrate ion is NO_3^-. We use one Ni^{2+} and two NO_3^- to make neutral $Ni(NO_3)_2$.

(b) Sodium ion is Na^+. Bicarbonate ion is HCO_3^-. We use one Na^+ and one HCO_3^- to make neutral $NaHCO_3$.

(c) Lithium ion is Li^+. Hypochlorite ion is ClO^-. We use one Li^+ and one ClO^- to make neutral LiClO.

(d) Magnesium ion is Mg^{2+}. Chlorate ion is ClO_3^-. We use one Mg^{2+} and two ClO_3^- to make neutral $Mg(ClO_3)_2$.

(e) Calcium ion is Mg^{2+}. Sulfite ion is SO_3^{2-}. We use one Mg^{2+} and one SO_3^{2-} to make neutral $CaSO_3$.

43. *Result:* **(a) $Fe(NO_3)_3$ (b) K_2CO_3 (c) Na_3PO_4 (d) $Ca(ClO_2)_2$ (e) Na_2SO_4**

Analyze, Plan, and Execute:

(a) Iron(III) ion is Fe^{3+}. Nitrate ion is NO_3^-. We use one Fe^{3+} and three NO_3^- to make neutral $Fe(NO_3)_3$.

(b) Potassium ion is K^+. Carbonate ion is CO_3^{2-}. We use two K^+ and one CO_3^{2-} to make neutral K_2CO_3.

(c) Sodium ion is Na^+. Phosphate ion is PO_4^{3-}. We use three Na^+ and one PO_4^{3-} to make neutral Na_3PO_4.

(d) Calcium ion is Ca^{2+}. Chlorite ion is ClO_2^-. We use one Mg^{2+} and two ClO_2^- to make neutral $Ca(ClO_2)_2$.

(e) Sodium ion is Na^+. Sulfate ion is SO_4^{2-}. We use two Na^+ and one SO_4^{2-} to make neutral Na_2SO_4.

44. *Result:* **(a) $(NH_4)_2CO_3$ (b) CaI_2 (c) $CuBr_2$ (d) $AlPO_4$**

Analyze and Plan: Make neutral combinations with the common ions involved.

Execute:

(a) Ammonium (NH_4^+) and carbonate (CO_3^{2-}) must be combined 2:1, to make $(NH_4)_2CO_3$.

(b) Calcium (Ca^{2+}) and iodide (I^-) must be combined 1:2, to make CaI_2.

(c) Copper(II) (Cu^{2+}) and bromide (Br^-) must be combined 1:2, to make $CuBr_2$.

(d) Aluminum (Al^{3+}) and phosphate (PO_4^{3-}) must be combined 1:1, to make $AlPO_4$.

45. *Result:* **(a) $Ca(HCO_3)_2$ (b) $KMnO_4$ (c) $Mg(ClO_4)_2$ (d) $(NH_4)_2HPO_4$**

Analyze and Plan: Make neutral combinations with the common ions involved.

Execute:

(a) Calcium (Ca^{2+}) and hydrogen carbonate (HCO_3^-) must be combined 1:2, to make $Ca(HCO_3)_2$.

(b) Potassium (K^+) and permanganate (MnO_4^-) must be combined 1:1, to make $KMnO_4$.

(c) Magnesium (Mg^{2+}) and perchlorate (ClO_4^-) must be combined 1:2, to make $Mg(ClO_4)_2$.

(d) Ammonium (NH_4^+) and monohydrogen phosphate (HPO_4^{2-}) must be combined 2:1, to make $(NH_4)_2HPO_4$.

46. *Result:* **(a) potassium sulfide (b) nickel(II) sulfate (c) ammonium phosphate (d) aluminum hydroxide (e) cobalt(III) sulfate**

Analyze and Plan: Give the name of the cation then the name of the anion.

Execute:

(a) K_2S contains cation K^+ called potassium and anion S^{2-} called sulfide, so it is potassium sulfide.

(b) $NiSO_4$ contains cation Ni^{2+} called nickel(II) and anion SO_4^{2-} called sulfate, so it is nickel(II) sulfate.

(c) $(NH_4)_3PO_4$ contains cation NH_4^+ called ammonium and anion PO_4^{3-} called phosphate, so it is ammonium phosphate.

(d) $Al(OH)_3$ contains cation Al^{3+} called aluminum and anion OH^- called hydroxide, so it is aluminum hydroxide.

(e) $Co_2(SO_4)_3$ contains cation Co^{3+} called cobalt(III) and anion SO_4^{2-} called sulfate, so it is cobalt(III) sulfate.

47. *Result:* **(a) potassium dihydrogen phosphate (b) copper(II) sulfate (c) chromium(III) chloride (d) calcium acetate (e) iron(III) sulfate**

Analyze and Plan: Give the name of the cation then the name of the anion.

Execute:

(a) KH_2PO_4 contains cation K^+ called potassium and anion $H_2PO_4^-$ called dihydrogen phosphate, so it is potassium dihydrogen phosphate.

(b) $CuSO_4$ contains cation Cu^{2+} called copper(II) and anion SO_4^{2-} called sulfate, so it is copper(II) sulfate.

(c) $CrCl_3$ contains cation Cr^{3+} called chromium(III) and anion Cl^- called chloride, so it is chromium(III) chloride.

(d) $Ca(CH_3COO)_2$ contains cation Ca^{2+} called calcium and anion CH_3COO^- called acetate, so it is calcium acetate.

(e) $Fe_2(SO_4)_3$ contains cation Fe^{3+} called iron(III) and anion SO_4^{2-} called sulfate, so it is iron(III) sulfate.

Ionic Compounds: Bonding and Properties (Section 2-6)

48. *Result:* **MgO; MgO has higher ionic charges and smaller ion sizes than NaCl**

Analyze, Plan, and Execute: Magnesium oxide is MgO, and it is composed of Mg^{2+} ions and O^{2-} ions. The relatively high melting temperature of MgO compared to NaCl (composed of Na^+ ions and Cl^- ions) is probably due to the larger ionic charges and smaller sizes of the ions. The large opposite charges sitting close together have very strong attractive forces between the ions. Melting requires that these attractive forces be overcome.

49. *Result:* **$NaNO_3$; ionic compound is solid, NO_2 and NH_3 are covalent gases.**

Analyze, Plan, and Execute: A white crystalline powder in a bottle that melts at 310 °C is probably the ionic compound, $NaNO_3$. NO_2 and NH_3 are covalent compounds and are in the gaseous state at room temperature.

Molecular Compounds (Section 2-7)

50. *Result:* (a) ionic (b) molecular (c) molecular (d) ionic

Analyze and Plan: To tell if a compound is ionic or not, look at the formula for metals and nonmetals together, or common cations and anions. If a compound contains only nonmetals or metalloids and nonmetals, it is probably molecular. Ionic compounds have very high melting points (well above room temperature) and will conduct electricity when melted.

Execute:

(a) Rb_2O has a metal and a nonmetal together. **Ionic**.

(b) C_6H_{12} contains only nonmetals. **Molecular**.

(c) A compound that is a liquid at room temperature. **Molecular**.

(d) A compound that conducts electricity when molten. **Ionic**.

51. *Result:* **(a) ionic (b) molecular (c) ionic (d) molecular**

Analyze and Plan: To tell if a compound is ionic or not, look at the formula for metals and nonmetals together, or common cations and anions. If a compound contains only nonmetals or metalloids and nonmetals, it is probably molecular. Ionic compounds have very high melting points (well above room temperature) and can be cleaved with a sharp wedge.

Execute:

(a) A compound that can be cleaved with a sharp wedge. **Ionic**.

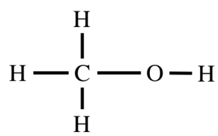
(b) A compound that is a liquid at room temperature. **Molecular**.

(c) $MgBr_2$ has a metal and a nonmetal together. **Ionic**.

(d) $C_5H_{10}O_2N$ contains only nonmetals. **Molecular**.

52. *Result/Explanation:*

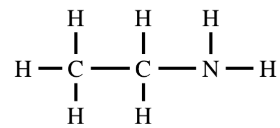
(a) Structural



Molecular:

CH_4O

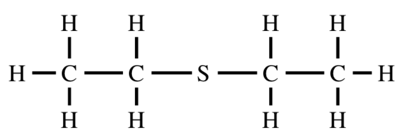
(b) Structural



Molecular:

C_2H_7N

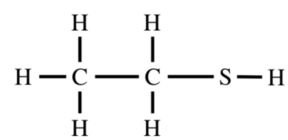
(c) Structural



Molecular:

$C_4H_{10}S$

(d) Structural

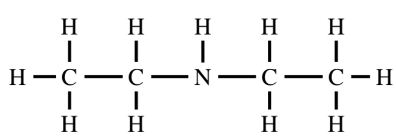


Molecular:

C_2H_6S

53 *Result/Explanation:*

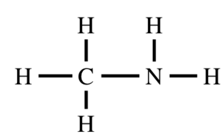
(a) Structural



Molecular:

$C_4H_{11}N$

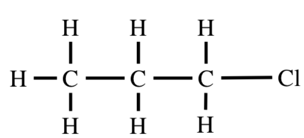
(b) Structural



Molecular:

CH_5N

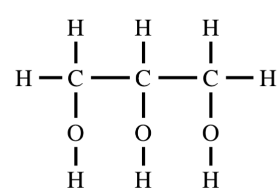
(c) Structural



Molecular:

C_3H_7Cl

(d) Structural



Molecular:

$C_3H_{10}O_3$

54. *Result/Explanation:*

(a) Heptane Molecular Formula: $\mathbf{C_7H_{16}}$

(b) Acrylonitrile Molecular Formula: $\mathbf{C_3H_3N}$

55. *Result/Explanation:*

(a) Fenclorac Molecular Formula: $\mathbf{C_{14}H_{16}Cl_2O_2}$

(b) Vitamin B012 Molecular Formula: $\mathbf{C_{63}H_{88}CoN_{14}O_{14}P}$

56. *Result:* **(a) 1 Ca, 2 C, 4 O (b) 8 C, 8 H (c) 2 N, 8 H, 1 S, 4 O (d) 1 Pt, 2 N, 6 H, 2 Cl (e) 4 K, 1 Fe, 6 C, 6 N**

Analyze and Plan: Keep in mind that atoms found inside parentheses that are followed by a subscript get multiplied by that subscript.

Execute:

(a) CaC_2O_4 contains one atom of calcium, two atoms of carbon, and four atoms of oxygen.

(b) $C_6H_5CHCH_2$ contains eight atoms of carbon and eight atoms of hydrogen.

(c) $(NH_4)_2SO_4$ contains two (1 × 2) atoms of nitrogen, eight (4 × 2) atoms of hydrogen, one atom of sulfur, and four atoms of oxygen.

(d) $Pt(NH_3)_2Cl_2$ contains one atom of platinum, two (1 × 2) atoms of nitrogen, six (3 × 2) atoms of hydrogen, and two atoms of chlorine.

(e) $K_4Fe(CN)_6$ contains four atoms of potassium, one atom of iron, six (1 × 6) atoms of carbon, and six (1 × 6) atoms of nitrogen.

57. *Result:* **(a) 9 C, 10 H, 2 O (b) 4 C, 4 O, 6 H (c) 1 N, 7 H, 3 C, 2 O (d) 10 C, 11 H, 1 N, 1 Fe (e) 7 C, 5 H, 3 N, 6 O**

Analyze and Plan: Keep in mind that atoms found inside parentheses that are followed by a subscript get multiplied by that subscript.

Execute:

(a) $C_6H_5COOC_2H_5$ contains nine atoms of carbon, ten atoms of hydrogen, and two atoms of oxygen.

(b) $HOOCCH_2CH_2COOH$ contains four atoms of carbon, four atoms of oxygen, and six atoms of hydrogen.

(c) $NH_2CH_2CH_2COOH$ contains one atom of nitrogen, seven (2 + 2 + 2 + 1) atoms of hydrogen, three (1 + 1 + 1) atoms of carbon, and two atoms of oxygen.

(d) $C_{10}H_9NH_2Fe$ contains ten atoms of carbon, eleven (9 + 2) atoms of hydrogen, one atom of nitrogen, and one atom of iron.

(e) $C_6H_2CH_3(NO_2)_3$ contains seven atoms of carbon, five atoms of hydrogen, three (1 × 3) atoms of nitrogen, six (2 × 3) atoms of oxygen.

Naming Binary Molecular Compounds (Section 2-8)

58. *Result/Explanation:* A general rule for naming binary compounds is to name the first element then take the first part of the name of the second element and add the ending -ide. Prefixes given in Table 2.6 are used to designate the number of a particular kind of atom, such as mono- for one, di- for two, tri- for three, etc.

(a) SO_2 is **sulfur dioxide**.

(b) CCl_4 is **carbon tetrachloride**.

(c) P_4S_{10} is **tetraphosphorus decasulfide**.

(d) SF_4 is **sulfur tetrafluoride**.

59. *Result/Explanation:* A general rule for naming binary compounds is to name the first element then take the first part of the name of the second element and add the ending -ide. Prefixes given in Table 2.6 are used to designate the number of a particular kind of atom, such as mono- for one, di- for two, tri- for three, etc.

(a) HBr is **hydrogen bromide**.

(b) ClF_3 is **chlorine trifluoride**.

(c) Cl_2O_7 is **dichlorine heptaoxide**.

(d) BI_3 is **boron triiodide**.

60. *Result/Explanation:* A general rule for applying the names of binary compounds to the formula is to list the symbol for first element named then the symbol for the second element. Use the prefixes described in Table 2.6 to learn the number of a particular kind of atom and use that number for the subscript on the symbol.

(a) nitrogen triiodide has an N atom and three I atoms: **NI_3**.

(b) carbon disulfide has a C atom and two S atoms. **CS_2**.

(c) dinitrogen tetraoxide has two N atoms and four O atoms: **N_2O_4**.

(d) selenium hexafluoride has one Se atom and six F atoms: **SeF_6**.

61. *Result/Explanation:* A general rule for applying the names of binary compounds to the formula is to list the symbol for first element then the symbol for the second element. Use the prefixes described in Table 2.6 to learn the number of a particular kind of atom and use that number for the subscript on the symbol.

(a) bromine triichloride has an Br atom and three Cl atoms: **$BrCl_3$**.

(b) xenon trioxide has a Xe atom and three O atoms. **XeO_3**.

(c) diphosphorus tetrafluoride has two P atoms and four F atoms: **P_2F_4**.

(d) oxygen difluoride has one O atom and two F atoms: **OF_2**.

Organic Molecular Compounds (Section 2-9)

62. *Result/Explanation:* Carbon makes four bonds. A carbon atom in an alkane chain is bonded to at least one other C atom, so that leaves up to three remaining bonds that may each be to an H atom. So, in a noncyclic alkane other than methane, the maximum number of hydrogen atoms that can be bonded to one carbon atom is **three**.

63. *Result/Explanation:* Carbon makes four bonds. A carbon atom in an alkane chain may be bonded to as many as **four** other C atoms.

64. *Result/Explanation:*

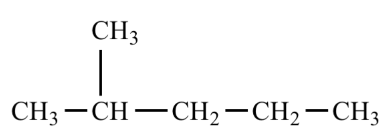
(a) Two molecules that are constitutional isomers have the same formula (i.e., on the molecular level, these molecules have **the same number of atoms of each kind**).

(b) Two molecules that are constitutional isomers of each other have their atoms in **different bonding arrangements**.

65. *Result/Explanation:* Five constitutional hexane isomers:

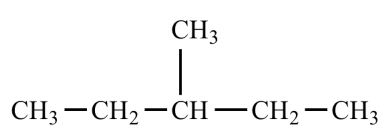
(1) Straight six-carbon chain: CH_3—CH_2—CH_2—CH_2—CH_2—CH_3

(2) Five-carbon chain with a branch on the second carbon:



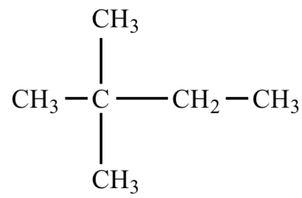
The condensed structural formula looks like this: $CH_3CH(CH_3)CH_2CH_2CH_3$

(3) Five-carbon chain with branch on the third carbon:



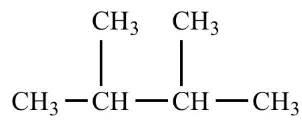
The condensed structural formula looks like this: $CH_3CH_2CH(CH_3)CH_2CH_3$

(4) Four-carbon chain with two branches on the second carbon:



The condensed structural formula looks like this: $CH_3C(CH_3)_2CH_2CH_3$

(5) Four-carbon chain with branches on the second carbon and a branch on the third carbon:



The condensed structural formula looks like this: $CH_3CH(CH_3)CH(CH_3)_2$

66. *Result/Explanation:* Noncyclic hydrocarbons have 2n + 2 hydrogen atoms, where n = number of carbon atoms. Eicosane has 20 carbon atoms, so it has 2(20) + 2 = **42 hydrogen atoms**.

67. *Result/Explanation:* Cyclic hydrocarbons have 2n hydrogen atoms, where n = number of carbon atoms. A cyclic hydrocarbon with 16 hydrogen atoms has 2n = 16 hydrogen atoms. n = **8 carbon atoms**. The cyclic hydrocarbon with eight carbons is called **cyclooctane**.

Amount of Substance: The Mole (Section 2-10)

68. *Result:* **2 × 10⁸ years**

Analyze: Determine how long it will take for all the people in the United States to count 1 mole of pennies if they spend eight hours a day every day counting.

Plan: Calculate the number of pennies each person has to count, then calculate how many days each person would spend counting their share.

Execute:

$$\frac{6.022\times10^{23}\ \text{pennies}}{300{,}000{,}000\ \text{people}} = 2\times10^{15}\ \text{pennies/person}$$

$$\frac{2\times10^{15}\ \text{pennies}}{\text{person}}\times\frac{1\,\text{s}}{1\,\text{penny}}\times\frac{1\,\text{min.}}{60\,\text{s}}\times\frac{1\,\text{hour}}{60\,\text{min.}}\times\frac{1\,\text{day}}{8\,\text{hours}}\times\frac{1\,\text{year}}{365.25\ \text{days}} = 2\times10^{8}\ \text{years}$$

Assuming that the population stays fixed over this period of time and that no one quits the job or dies without being replaced, it would take about 200 billion years for the people in the United States to count this one mole of pennies.

✓ *Reasonable Result Check:* The quantity of pennies in one mole is huge. It will take people a LONG time to count that many pennies.

69. *Result/Explanation:* Counting the individual molecules is inconvenient for two reasons. Individual molecules are too small, and in samples large enough to see, their numbers are so great that not even normal "large number" words are very convenient. For example, a common "large number" word is "trillion". That's 1,000,000,000,000 or 1×10^{12}. One mole of molecules is almost a trillion times more than a trillion!

Molar Mass (Section 2-11)

70. *Result:* **(a) 27 g B (b) 0.48 g O_2 (c) 6.98×10^{-2} g Fe (d) 2.61×10^{3} g He**

Analyze: Determine mass in grams from given quantity in moles.

Plan: Look up the elements on the periodic table to get the atomic weight (with at least four significant figures). If necessary, calculate the molecular weight. Use that number for the molar mass (with units of grams per mole) as a conversion factor between moles and grams.

Notice: Whenever you use physical constants that you look up, it is important to carry more significant figures than the rest of the measured numbers, to prevent causing inappropriate round-off errors.

Execute:

(a) Boron (B) has atomic number 5 on the periodic table. Its atomic weight is 10.811 u/atom, so the molar mass is 10.811 g/mol.

$$2.5 \text{ mol B} \times \frac{10.811 \text{ g B}}{1 \text{ mol B}} = 27 \text{ g B}$$

(b) O_2 (diatomic molecular oxygen) is made with two atoms of the element with the atomic number 8 on the periodic table. Its atomic weight is 15.9994 u/atom; therefore, the molecular weight of O_2 is 2×15.9994 u/atom = 31.9988 u/molecule, and the molar mass is 31.9988 g/mol.

$$0.015 \text{ mol } O_2 \times \frac{31.9988 \text{ g } O_2}{1 \text{ mol } O_2} = 0.48 \text{ g } O_2$$

(c) Iron (Fe) has atomic number 26 on the periodic table. Its atomic weight is 55.845 u/atom, so the molar mass is 55.845 g/mol.

$$1.25 \times 10^{-3} \text{ mol Fe} \times \frac{55.845 \text{ g Fe}}{1 \text{ mol Fe}} = 6.98 \times 10^{-2} \text{ g Fe}$$

(d) Helium (He) has atomic number 2 on the periodic table. Its atomic weight is 4.0026 u/atom, so the molar mass is 4.0026 g/mol.

$$653 \text{ mol He} \times \frac{4.0026 \text{ g He}}{1 \text{ mol He}} = 2.61 \times 10^{3} \text{ g He}$$

☑ *Reasonable Result Check:* The mol units cancel when the factor is multiplied, leaving grams.

71. *Result:* **(a) 1.19 × 10^3 g Au (b) 11 g U (c) 315 g Ne (d) 0.0886 g Pu**

Analyze: Determine the mass in grams from given quantity in moles.

Plan: Look up the elements on the periodic table to get the atomic weight (with at least four significant figures). Use that number for the molar mass (with units of grams per mole) as a conversion factor between moles and grams.

Notice: Whenever you use physical constants that you look up, it is important to carry more significant figures than the rest of the measured numbers, to prevent causing inappropriate round-off errors.

Execute:

(a) Gold (Au) has atomic number 79 on the periodic table. Its atomic weight is 196.9666 u/atom, so the molar mass is 196.9666 g/mol.

$$6.03 \text{ mol Au} \times \frac{196.9666 \text{ g Au}}{1 \text{ mol Au}} = 1.19 \times 10^{3} \text{ g Au}$$

(b) Uranium (U) has atomic number 92 on the periodic table. Its atomic weight is 238.0289 u/atom, so the molar mass is 238.0289 g/mol.

$$0.045 \text{ mol U} \times \frac{238.0289 \text{ g U}}{1 \text{ mol U}} = 11 \text{ g U}$$

(c) Neon (Ne) has atomic number 10 on the periodic table. Its atomic weight is 20.1797 u/atom, so the molar mass is 20.1797 g/mol.

$$15.6 \text{ mol Ne} \times \frac{20.1797 \text{ g Ne}}{1 \text{ mol Ne}} = 315 \text{ g Ne}$$

(d) Radioactive plutonium (Pu) has atomic number 94 on the periodic table. The atomic weight given on the periodic table is the weight of its most stable isotope 244 u/atom, so the molar mass is 244 g/mol.

$$3.63\times10^{-4}\text{ mol Pu}\times\frac{244\text{ g Pu}}{1\text{ mol Pu}}=0.0886\text{ g Pu}$$

☑ *Reasonable Result Check:* The mol units cancel when the factor is multiplied, leaving the answer in grams.

72. *Result:* **(a) 1.9998 mol Cu (b) 0.499 mol Ca (c) 0.6208 mol Al (d) 3.1 × 10^{-4} mol K (e) 2.1 × 10^{-5} mol Am**

Analyze: Determine the quantity in moles from given mass in grams.

Plan: Look up the elements on the periodic table to get the atomic weight. Use that number for the molar mass (with units of grams per mole) as a conversion factor between grams and moles.

Notice: Whenever you use physical constants that you look up, it is important to carry more significant figures than the rest of the measured numbers, to prevent causing inappropriate round-off errors.

Execute:

(a) Copper (Cu) has atomic number 29 on the periodic table. Its atomic weight is 63.546 u/atom, so the molar mass is 63.546 g/mol.

$$127.08\text{ g Cu}\times\frac{1\text{ mol Cu}}{63.546\text{ g Cu}}=1.9998\text{ mol Cu}$$

(b) Calcium (Ca) has atomic number 20 on the periodic table. Its atomic weight is 40.078 u/atom, so the molar mass is 40.078 g/mol.

$$20.0\text{ g Ca}\times\frac{1\text{ mol Ca}}{40.078\text{ g Ca}}=0.499\text{ mol Ca}$$

(c) Aluminum (Al) has atomic number 13 on the periodic table. Its atomic weight is 26.9815 u/atom, so the molar mass is 26.9815 g/mol.

$$16.75\text{ g Al}\times\frac{1\text{ mol Al}}{26.9815\text{ g Al}}=0.6208\text{ mol Al}$$

(d) Potassium (K) has atomic number 19 on the periodic table. Its atomic weight is 39.0983 u/atom, so the molar mass is 39.0983 g/mol.

$$0.012\text{ g K}\times\frac{1\text{ mol K}}{39.0983\text{ g K}}=3.1\times10^{-4}\text{ mol K}$$

(e) Radioactive americium (Am) has atomic number 95 on the periodic table. The atomic weight given on the periodic table is the weight of its most stable isotope 243 u/atom, so the molar mass is 243 g/mol.

Convert milligrams into grams, first.

$$5.0\text{ mg Am}\times\frac{1\text{ g Am}}{1000\text{ mg Am}}\times\frac{1\text{ mol Am}}{243\text{ g Am}}=2.1\times10^{-5}\text{ mol Am}$$

☑ *Reasonable Result Check:* Notice that grams units cancel when the factor is multiplied, leaving moles.

73. *Result:* **(a) 0.696 mol Na (b) 1.7 × 10^{-5} mol Pt (c) 0.0497 mol P (d) 0.0117 mol As**

(e) 7.49 × 10^{-3} mol Xe

Analyze: Determine the quantity in moles from given mass in grams.

Plan: Look up the elements on the periodic table to get the atomic weight. Use that number for the molar mass (with units of grams per mole) as a conversion factor between grams and moles.

Execute:

(a) Sodium (Na) has atomic number 11 on the periodic table. Its atomic weight is 22.9898 u/atom, so the molar mass is 22.9898 g/mol.

$$16.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.9898 \text{ g Na}} = 0.696 \text{ mol Na}$$

(b) Platinum (Pt) has atomic number 78 on the periodic table. Its atomic weight is 195.078 u/atom, so the molar mass is 195.078 g/mol.

$$0.0034 \text{ g Pt} \times \frac{1 \text{ mol Pt}}{195.078 \text{ g Pt}} = 1.7 \times 10^{-5} \text{ mol Pt}$$

(c) Phosphorus (P) has atomic number 15 on the periodic table. Its atomic weight is 30.9738 u/atom, so the molar mass is 30.9738 g/mol.

$$1.54 \text{ g P} \times \frac{1 \text{ mol P}}{30.9738 \text{ g P}} = 0.0497 \text{ mol P}$$

(d) Arsenic (As) has atomic number 33 on the periodic table. Its atomic weight is 74.9216 u/atom, so the molar mass is 74.9216 g/mol.

$$0.876 \text{ g As} \times \frac{1 \text{ mol As}}{74.9216 \text{ g As}} = 0.0117 \text{ mol As}$$

(e) Xenon (Xe) has atomic number 54 on the periodic table. Its atomic weight is 131.29 u/atom, so the molar mass is 131.29 g/mol.

$$0.983 \text{ g Xe} \times \frac{1 \text{ mol Xe}}{131.29 \text{ g Xe}} = 7.49 \times 10^{-3} \text{ mol Xe}$$

☑ *Reasonable Result Check:* Notice that grams units cancel when the factor is multiplied, leaving just moles.

74. *Result:* **4.131×10^{23} Cr atoms**

Analyze: Given a chromium sample with known mass, determine the number of atoms in it.

Plan: Start with the mass. Use the molar mass of chromium as a conversion factor between grams and moles. Then use Avogadro's number as a conversion factor between moles of chromium atoms and the actual number of chromium atoms.

Execute:

$$35.67 \text{ g Cr} \times \frac{1 \text{ mol Cr atoms}}{51.996 \text{ g Cr}} \times \frac{6.0221 \times 10^{23} \text{ Cr atoms}}{1 \text{ mol Cr atoms}} = 4.131 \times 10^{23} \text{ Cr atoms}$$

☑ *Reasonable Result Check:* A sample of chromium that a person can see and hold is macroscopic. It will contain a very large number of atoms.

75. *Result:* **5.93×10^{21} Au atoms**

Analyze: A ring of gold has a known mass. Determine the number of atoms in the sample.

Plan: Start with the mass. Use the molar mass of gold as a conversion factor between grams and moles. Then use Avogadro's number as a conversion factor between moles of gold atoms and the number of gold atoms.

Execute:

$$1.94 \text{ g Au} \times \frac{1 \text{ mol Au atoms}}{196.9666 \text{ g Au}} \times \frac{6.022 \times 10^{23} \text{ Au atoms}}{1 \text{ mol Au atoms}} = 5.93 \times 10^{21} \text{ Au atoms}$$

☑ *Reasonable Result Check:* A ring of gold is something that a person can see and hold. It will contain a very large number of atoms.

76. *Result:* **(a) 12.63 g (b) 7.689×10^{20} molecules (c) 1.538×10^{21} N atoms (d) 14.4 g N**

Analyze: $C_{13}H_{10}N_2$ has 13 C atoms, 10 H atoms, and 2 N atoms.

Plan: Look up the elements on the periodic table to get their atomic weights. Combine those numbers for the molar mass (with units of grams per mole):

Execute: Carbon (C), with atomic number 6, has a molar mass of 12.0107 g/mol. Hydrogen (H), with atomic number 1, has a molar mass of 1.0079 g/mol. Nitrogen, with atomic number 7, has a molar mass of 14.0067 g/mol.

$$\left(\frac{13 \text{ mol C}}{1 \text{ mol comp}}\right)\times\left(\frac{12.0107\text{g}}{1 \text{ mol C}}\right)+\left(\frac{10 \text{ mol H}}{1 \text{ mol comp}}\right)\times\left(\frac{1.0079 \text{ g}}{1 \text{ mol H}}\right)+\left(\frac{2 \text{ mol N}}{1 \text{ mol comp}}\right)\times\left(\frac{14.0067 \text{ g}}{1 \text{ mol N}}\right)= 194.2315 \text{ g/mol}$$

A shorthand version of this calculation looks like this:

13(12.0107 g/mol C) + 10(1.0079 g/mol H) + 2(14.0067 g/mol N) = 194.2315 g/mol $C_{13}H_{10}N_2$

(a) Use molar mass as a conversion factor between moles and grams.

$$0.06500 \text{ mol comp} \times \frac{194.2315 \text{ g comp}}{1 \text{ mol comp}} = 12.63 \text{ g comp}$$

(b) Use Avogadro's number to relate moles to molecules.

$$0.2480 \text{ g comp} \times \left(\frac{1 \text{ mol comp}}{194.2315 \text{ g comp}}\right) \times \left(\frac{6.022 \times 10^{23} \text{ molecules comp}}{1 \text{ mol comp}}\right) = 7.689 \times 10^{20} \text{ molecules comp}$$

(c) Use the chemical formula, $C_{13}H_{10}N_2$, to relate the atoms of N to molecules of compound.

$$7.689 \times 10^{20} \text{ molecules comp} \times \left(\frac{2 \text{ N atoms}}{1 \text{ molecule comp}}\right) = 1.538 \times 10^{21} \text{ N atoms}$$

(d) Calculate moles of compound using the molar mass, use the formula to relate moles of N atoms, then use the molar mass of N to calculate mass of nitrogen.

$$100. \text{ g comp} \times \left(\frac{1 \text{ mol comp}}{194.2315 \text{ g comp}}\right) \times \left(\frac{2 \text{ mol N}}{1 \text{ mol comp}}\right) \times \left(\frac{14.0067 \text{ g N}}{1 \text{ mol N}}\right) = 114.4 \text{ g N}$$

☑ *Reasonable Result Check:* Notice that several units cancel when the factors are multiplied.

77. *Result:* **(a) 12.63 g (b) 7.689 × 10^{20} molecules (c) 1.538 × 10^{21} N atoms (d) 14.4 g N**

Analyze, Plan, and Execute: $C_{12}H_{24}N_9P_3$ has 12 C atoms, 24 H atoms, 9 N atoms, and 3 P.

(a) Look up the elements on the periodic table to get their atomic weights. Combine those numbers for the molar mass (with units of grams per mole):

Carbon (C), with atomic number 6, has a molar mass of 12.0107 g/mol. Hydrogen (H), with atomic number 1, has a molar mass of 1.0079 g/mol. Nitrogen, with atomic number 7, has a molar mass of 14.0067 g/mol. Phosphorus, with atomic number 15, has a molar mass of 30.9738 g/mol.

$$\left(\frac{12 \text{ mol C}}{1 \text{ mol comp}}\right)\times\left(\frac{12.0107\text{g}}{1 \text{ mol C}}\right)+\left(\frac{24 \text{ mol H}}{1 \text{ mol comp}}\right)\times\left(\frac{1.0079 \text{ g}}{1 \text{ mol H}}\right)+$$
$$\left(\frac{9 \text{ mol N}}{1 \text{ mol comp}}\right)\times\left(\frac{14.0067 \text{ g}}{1 \text{ mol N}}\right)+\left(\frac{3 \text{ mol P}}{1 \text{ mol comp}}\right)\times\left(\frac{30.9738 \text{ g}}{1 \text{ mol P}}\right)= 387.2997 \text{ g/mol}$$

A shorthand version of this calculation looks like this:

12(12.0107 g/mol C) + 24(1.0079 g/mol H) + 9(14.0067 g/mol N)

+ 3(30.9738 g/mol P) = 387.2997 g/mol $C_{12}H_{24}N_9P_3$

(b) Calculate moles of compound using molar mass, use the formula to relate moles of N atoms, then use the molar mass of N to calculate mass of nitrogen.

$$100. \text{ g comp} \times \left(\frac{1 \text{ mol comp}}{387.2997 \text{ g comp}}\right) \times \left(\frac{9 \text{ mol N}}{1 \text{ mol comp}}\right) \times \left(\frac{14.0067 \text{ g N}}{1 \text{ mol N}}\right) = 32.5 \text{ g N}$$

(c) Calculate moles of compound using molar mass, use the formula to relate moles of P atoms, then use the molar mass of P to calculate mass of phosphorus.

$$250.0 \text{ mg comp} \times \left(\frac{1 \text{ g}}{1000 \text{ mg}}\right) \times \left(\frac{1 \text{ mol comp}}{387.2997 \text{ g comp}}\right) \times \left(\frac{3 \text{ mol P}}{1 \text{ mol comp}}\right) = 1.936 \times 10^{-3} \text{ mol P}$$

$$1.936 \times 10^{-3} \text{ mol P} \times \left(\frac{30.9738 \text{ g P}}{1 \text{ mol P}}\right) = 0.05998 \text{ g P}$$

(d) Use Avogadro's number to relate moles to atoms:

$$1.936 \times 10^{-3} \text{ mol P} \times \left(\frac{6.022 \times 10^{23} \text{ P atoms}}{1 \text{ mol P}}\right) = 1.166 \times 10^{21} \text{ P atoms}$$

☑ *Reasonable Result Check:* Notice that several units cancel when the factors are multiplied.

78. *Result:* **(a) 41.7 pennies (b) 9.28 × 10 $^{-4}$ mol Cu (c) 5.59 × 10^{20} Cu atoms**

Analyze, Plan, and Execute:

(a) Use the percent Cu in a penny and the mass of one penny to calculate the number of pennies.

$$2.458 \text{ g Cu} \times \left(\frac{100 \text{ g penny}}{2.40 \text{ g Cu}}\right) \times \left(\frac{1 \text{ g penny}}{2.458 \text{ g penny}}\right) = 41.7 \text{ pennies}$$

(b) Calculate moles of copper using the molar mass of Cu, 63.546 g/mol.

$$2.458 \text{ g penny} \times \left(\frac{2.40 \text{ g Cu}}{100 \text{ g penny}}\right) \times \left(\frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}}\right) = 9.28 \times 10^{-4} \text{ mol Cu}$$

(c) Use Avogadro's number to relate moles to atoms.

$$9.28 \times 10^{-4} \text{ mol Cu} \times \left(\frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}}\right) = 5.59 \times 10^{20} \text{ Cu atoms}$$

☑ *Reasonable Result Check:* Notice that several units cancel when the factors are multiplied.

79. *Result:* **(a) 40.0 g Ag (b) 0.370 mol Ag (c) 2.23 × 10^{23} Ag atoms**

Analyze, Plan, and Execute:

(a) Use the percent Ag in sterling silver. $43.2 \text{ g sterling} \times \left(\frac{92.5 \text{ g Ag}}{100 \text{ g sterling}}\right) = 40.0 \text{ g Ag}$

(b) Calculate moles of silver using the molar mass of Ag, 107.8682 g/mol.

$$40.0 \text{ g Ag} \times \left(\frac{1 \text{ mol Ag}}{107.8682 \text{ g Ag}}\right) = 0.370 \text{ mol Ag}$$

(c) Use Avogadro's number to relate moles to atoms.

$$0.370 \text{ mol Ag} \times \left(\frac{6.022 \times 10^{23} \text{ Ag atoms}}{1 \text{ mol Ag}}\right) = 2.23 \times 10^{23} \text{ Ag atoms}$$

☑ *Reasonable Result Check:* Notice that several units cancel when the factors are multiplied.

80. *Result:*

	CH_3OH	Carbon	Hydrogen	Oxygen
Amount of substance (mol)	**1 mol**	**1 mol**	**4 mol**	**1 mol**
No. of molecules or atoms	**6.022×10^{23} molecules**	**6.022×10^{23} atoms**	**2.409×10^{24} atoms**	**6.022×10^{23} atoms**
Molar mass (grams per mol methanol)	**32.0417 g/mol**	**12.0107 g/mol**	**4.0316 g/mol**	**15.9994 g/mol**

Analyze, Plan, and Execute: Consider a sample of 1 mol of methanol. The formula gives the mole ratio of each atom in one mole of methanol. Each mole of atoms is 6.022×10^{23} atoms. The molar masses can be determined by looking up each element on the period table, finding the atomic weight (which is also the grams/mol), and multiplying by the number of moles (see the solution to Question 77 for more details).

81. *Result:*

	$C_6H_{12}O_6$	Carbon	Hydrogen	Oxygen
Amount of substance (mol)	**1 mol**	**6 mol**	**12 mol**	**6 mol**
No. of molecules or atoms	**6.022×10^{23} molecules**	**3.613×10^{24} atoms**	**7.226×10^{24} atoms**	**3.613×10^{24} atoms**
Molar mass (grams per mol glucose)	**180.1554 g/mol**	**72.0642 g/mol**	**12.0948 g/mol**	**95.9964 g/mol**

Analyze, Plan, and Execute: Consider a sample of 1 mol of glucose. The formula gives the mole ratio of each atom in one mole of glucose. Each mole is 6.022×10^{23} molecules. The molar masses can be determined by looking up each element on the period table, finding the atomic weight (which is also the grams/mol), and multiplying by the number of moles.

82. *Result:* **(a) 0.0312 mol (b) 0.0101 mol (c) 0.0125 mol (d) 0.00406 mol (e) 0.00599 mol**

Analyze: Determine the amount (in moles) in a given mass of a compound.

Plan: Use the formula and the periodic table to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute:

(a) Molar mass CH_3OH = (12.0107 g/mol C) + 4(1.0079 g/mol H)

+ (15.9994 g/mol O) = 32.0417 g/mol CH_3OH

$$1.00 \text{ g } CH_3OH \times \frac{1 \text{ mol } CH_3OH}{32.0417 \text{ g } CH_3OH} = 0.0312 \text{ mol } CH_3OH$$

(b) Molar mass Cl_2CO = 2(35.453 g/mol Cl) + (12.0107 g/mol C) + (15.9994 g/mol O) = 98.916 g/mol Cl_2CO

$$1.00 \text{ g } Cl_2CO \times \frac{1 \text{ mol } Cl_2CO}{98.916 \text{ g } Cl_2CO} = 0.0101 \text{ mol } Cl_2CO$$

(c) Ammonium nitrate is NH_4NO_3.
Molar mass NH_4NO_3 = 2(14.0067 g/mol N) + 4(1.0079 g/mol H)

+ 3(15.9994 g/mol O) = 80.0432 g/mol NH_4NO_3

$$1.00 \text{ g } NH_4NO_3 \times \frac{1 \text{ mol } NH_4NO_3}{80.0432 \text{ g } NH_4NO_3} = 0.0125 \text{ mol } NH_4NO_3$$

(d) Magnesium sulfate heptahydrate is $MgSO_4 \cdot 7\ H_2O$.

Molar mass $MgSO_4 \cdot 7\ H_2O$ = (24.305 g/mol Mg) + (32.065 g/mol S)

+ 11(15.9994 g/mol O) + 14(1.0079 g/mol H) = 246.474 g/mol $MgSO_4 \cdot 7\ H_2O$

$$1.00 \text{ g } MgSO_4 \cdot 7H_2O \times \frac{1 \text{ mol } MgSO_4 \cdot 7H_2O}{246.474 \text{ g } MgSO_4 \cdot 7H_2O} = 0.00406 \text{ mol } MgSO_4 \cdot 7H_2O$$

(e) Silver acetate is $AgC_2H_3O_2$.

Molar mass $AgC_2H_3O_2$ = (107.8682 g/mol Ag) + 2(12.0107 g/mol C)

+ 3(1.0079 g/mol H) + 2(15.9994 g/mol O) = 166.9121 g/mol $AgC_2H_3O_2$

$$1.00 \text{ g } AgC_2H_3O_2 \times \frac{1 \text{ mol } AgC_2H_3O_2}{166.9121 \text{ g } AgC_2H_3O_2} = 0.00599 \text{ mol } AgC_2H_3O_2$$

☑ *Reasonable Result Check:* The quantity in moles is always going to be smaller than the mass in grams.

83. *Result:* **(a) 0.00136 mol (b) 0.00106 mol (c) 7.85 × 10^{-4} mol**

Analyze: Determine the amount (in moles) in a given mass of a compound.

Plan: Use the formula and the periodic table to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute:

(a) Molar mass $C_7H_5NO_3S$ = 7(12.0107 g/mol C) + 5(1.0079 g/mol H) + (14.0067 g/mol N)

+ 3(15.9994 g/mol O) + (32.065 g/mol S) = 183.184 g/mol $C_7H_5NO_3S$

$$0.250 \text{ g } C_7H_5NO_3S \times \frac{1 \text{ mol } C_7H_5NO_3S}{183.184 \text{ g } C_7H_5NO_3S} = 0.00136 \text{ mol } C_7H_5NO_3S$$

(b) Molar mass $C_{13}H_{20}N_2O_2$ = 13(12.0107 g/mol C) + 20(1.0079 g/mol H)

+ 2(14.0067 g/mol N) + 2(15.9994 g/mol O) = 236.3093 g/mol $C_{13}H_{20}N_2O_2$

$$0.250 \text{ g } C_{13}H_{20}N_2O_2 \times \frac{1 \text{ mol } C_{13}H_{20}N_2O_2}{236.3093 \text{ g } C_{13}H_{20}N_2O_2} = 0.00106 \text{ mol } C_{13}H_{20}N_2O_2$$

(c) Molar mass $C_{20}H_{14}O_4$ = 20(12.0107 g/mol C) + 14(1.0079 g/mol H)

+ 4(15.9994 g/mol O) = 318.3222 g/mol $C_{20}H_{14}O_4$

$$0.250 \text{ g } C_{20}H_{14}O_4 \times \frac{1 \text{ mol } C_{20}H_{14}O_4}{318.3222 \text{ g } C_{20}H_{14}O_4} = 7.85 \times 10^{-4} \text{ mol } C_{20}H_{14}O_4$$

☑ *Reasonable Result Check:* The quantity in moles is always going to be smaller than the mass in grams.

84. *Result:* **(a) 151.1622 g/mol (b) 0.0352 mol (c) 25.1 g**

Analyze: Determine the molar mass of a compound and then determine the mass of a given number of moles and the number of moles in a given mass.

Plan: Use the formula and the periodic table to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute:

(a) Molar mass $C_8H_9O_2N$ = 8(12.0107 g/mol C) + 9(1.0079 g/mol H)

+ 2(15.9994 g/mol O) + (14.0067 g/mol N) = 151.1622 g/mol $C_8H_9O_2N$

(b)
$$5.32 \text{ g } C_8H_9O_2N \times \frac{1 \text{ mol } C_8H_9O_2N}{151.1622 \text{ g } C_8H_9O_2N} = 0.0352 \text{ mol } C_8H_9O_2N$$

(c)
$$0.166 \text{ mol } C_8H_9O_2N \times \frac{151.1622 \text{ g } C_8H_9O_2N}{1 \text{ mol } C_8H_9O_2N} = 25.1 \text{ g } C_8H_9O_2N$$

☑ *Reasonable Result Check:* The quantity in moles is always going to be smaller than the mass in grams.

85. *Result:* **(a) 0.00180 mol $C_9H_8O_4$, 0.02266 mol $NaHCO_3$, 0.005205 mol $C_6H_8O_7$**

(b) 1.08 × 10^{21} molecules

Analyze: Given the masses of three compounds in a mixture, determine the number of moles of each, then determine the number of molecules of one compound.

Plan: Use the formula and the periodic table to calculate the molar mass for the compound. Convert milligrams to grams, then use the molar mass as a conversion factor between grams and moles. Use Avogadro's number to determine the actual number of molecules.

Execute:

(a) Molar mass $C_9H_8O_4$ = 9(12.0107 g/mol C) + 8(1.0079 g/mol H)

+ 4(15.9994 g/mol O) = 180.1571 g/mol $C_9H_8O_4$

Molar mass $NaHCO_3$ = (22.98977 g/mol Na) + (1.0079 g/mol H)

+ (12.0107 g/mol C) + 3(15.9994 g/mol O) = 84.0066 g/mol $NaHCO_3$

Molar mass $C_6H_8O_7$ = 6(12.0107 g/mol C) + 8(1.0079 g/mol H)

+ 7(15.9994 g/mol O) = 192.1232 g/mol $C_6H_8O_7$

$$324\ \text{mg}\,C_9H_8O_4 \times \frac{1\ \text{g}\ C_9H_8O_4}{1000\ \text{mg}\ C_9H_8O_4} \times \frac{1\ \text{mol}\ C_9H_8O_4}{180.1571\ \text{g}\ C_9H_8O_4} = 0.00180\ \text{mol}\ C_9H_8O_4$$

$$1904\ \text{mg}\ NaHCO_3 \times \frac{1\ \text{g}\ NaHCO_3}{1000\ \text{mg}\ NaHCO_3} \times \frac{1\ \text{mol}\ NaHCO_3}{84.0066\ \text{g}\ NaHCO_3} = 0.02266\ \text{mol}\ NaHCO_3$$

$$1000.\ \text{mg}\ C_6H_8O_7 \times \frac{1\ \text{g}\ C_6H_8O_7}{1000\ \text{mg}\ C_6H_8O_7} \times \frac{1\ \text{mol}\ C_6H_8O_7}{192.1232\ \text{g}\ C_6H_8O_7} = 0.005205\ \text{mol}\ C_6H_8O_7$$

(b) $$0.00180\ \text{mol}\ C_9H_8O_4 \times \frac{6.022 \times 10^{23}\ C_9H_8O_4\ \text{molecules}}{1\ \text{mol}\ C_9H_8O_4} = 1.08 \times 10^{21}\ C_9H_8O_4\ \text{molecules}$$

☑ *Reasonable Result Check:* The quantity in moles is always going to be smaller than the mass in grams or milligrams. The number of molecules for a macroscopic sample will be huge.

86. *Result:* **2 × 10^{21} molecules**

Analyze: Given the volume of a compound and its density, determine the number of molecules of the compound.

Plan: Use the formula and the periodic table to calculate the molar mass for the compound. Use the density to convert the volume from milliliters to grams, then use the molar mass as a conversion factor between grams and moles, then use Avogadro's number to determine the number of molecules.

Execute: Molar mass H_2O = 2(1.0079 g/mol H) + (15.9994 g/mol O) = 18.0152 g/mol H_2O

$$\frac{1}{20}\ \text{mL}\ H_2O \times \frac{1.0\ \text{g}\ H_2O}{1\ \text{mL}\ H_2O} \times \frac{1\ \text{mol}\ H_2O}{18.0152\ \text{g}\ H_2O} \times \frac{6.022 \times 10^{23}\ H_2O\ \text{molecules}}{1\ \text{mol}\ H_2O} = 2 \times 10^{21}\ H_2O\ \text{molecules}$$

☑ *Reasonable Result Check:* The number of atoms in a macroscopic sample will be huge.

Composition and Chemical Formulas (Section 2-12)

87. *Result:* **(a) 239.3 g/mol PbS, 86.60% Pb, 13.40% S (b) 30.0688 g/mol C_2H_6, 79.8881% C, 20.1119% H (c) 60.0518 g/mol CH_3COOH, 40.0011% C, 6.7135% H, 53.2854% O (d) 80.0432 g/mol NH_4NO_3, 34.9979% C, 5.0368% H, 59.9654% O**

Analyze: Given the formula of a compound, determine the molar mass, and the mass percent of each element.

Plan: Calculate the mass of each element in one mole of compound, while calculating the molar mass of the compound. Divide the calculated mass of the element by the molar mass of the compound and multiply by 100% to get mass percent. To get the last element's mass percent, subtract the other percentages from 100%.

Execute: (a) Mass of Pb per mole of PbS = 207.2 g/mol Pb

Mass of S per mole of PbS = 32.065 g/mol S

Molar mass PbS = (207.2 g/mol Pb) + (32.065 g/mol S) = 239.3 g/mol PbS

$$\% \text{ Pb} = \frac{\text{mass of Pb per mol PbS}}{\text{mass of PbS per mol PbS}} \times 100\% = \frac{207.2 \text{ g Pb}}{239.3 \text{ g PbS}} \times 100\% = 86.60\% \text{ Pb in PbS}$$

$$\% \text{ S} = 100\% - 86.60\% \text{ Pb} = 13.40\% \text{ S in PbS}$$

(b) Mass of C per mole of C_2H_6 = 2(12.0107 g/mol C) = 24.0214 g/mol C

Mass of H per mole of C_2H_6 = 6(1.0079 g/mol H) = 6.0474 g/mol H

Molar mass C_2H_6 = (24.0214 g/mol C) + (6.0474 g/mol H) = 30.0688 g/mol C_2H_6

$$\% \text{ C} = \frac{\text{mass of C / mol } C_2H_6}{\text{mass of } C_2H_6 \text{ / mol } C_2H_6} \times 100\% = \frac{24.0214 \text{ g C}}{30.0688 \text{ g } C_2H_6} \times 100\% = 79.8881\% \text{ C in } C_2H_6$$

$$\% \text{ H} = 100\% - 79.8881\% \text{ C} = 20.1119\% \text{ H in } C_2H_6$$

(c) Mass of C per mole of CH_3COOH = 2(12.0107 g/mol C) = 24.0214 g/mol C

Mass of H per mole of CH_3COOH = 4(1.0079 g/mol H) = 4.0316 g/mol H

Mass of O per mole of CH_3COOH = 2(15.9994 g/mol O) = 31.9988 g/mol O

Molar mass CH_3COOH = (24.0214 g/mol C) + (4.0316 g/mol H) + (31.9988 g/mol O)

= 60.0518 g/mol CH_3COOH

$$\% \text{C} = \frac{\text{mass of C / mol } CH_3COOH}{\text{mass of } CH_3COOH \text{ / mol } CH_3COOH} \times 100\%$$

$$= \frac{24.0214 \text{ g C}}{60.0518 \text{ g } CH_3COOH} \times 100\% = 40.0011\% \text{ C in } CH_3COOH$$

$$\% \text{ H} = \frac{\text{mass of H / mol } CH_3COOH}{\text{mass of } CH_3COOH \text{ / mol } CH_3COOH} \times 100\%$$

$$= \frac{4.0316 \text{ g H}}{60.0518 \text{ g } CH_3COOH} \times 100\% = 6.7135\% \text{ H in } CH_3COOH$$

% O = 100% – 40.0011% C – 6.7135% H = 53.2854% O in CH_3COOH

(d) Mass of N per mole of NH_4NO_3 = 2(14.0067 g/mol N) = 28.0134 g/mol N

Mass of H per mole of NH_4NO_3 = 4(1.0079 g/mol H) = 4.0316 g/mol H

Mass of O per mole of NH_4NO_3 = 3(15.9994 g/mol O) = 47.9982 g/mol O

Molar mass NH_4NO_3 = (28.0134 g/mol N) + (4.0316 g/mol H)

+ (47.9982 g/mol O) = 80.0432 g/mol NH_4NO_3

$$\% \text{ N} = \frac{\text{mass of N / mol } NH_4NO_3}{\text{mass of } NH_4NO_3 \text{ / mol } NH_4NO_3} \times 100\% = \frac{28.0134 \text{ g N}}{80.0432 \text{ g } NH_4NO_3} \times 100\%$$

$$= 34.9979\% \text{ N in } NH_4NO_3$$

$$\% \text{ H} = \frac{\text{mass of H / mol } NH_4NO_3}{\text{mass of } NH_4NO_3 \text{ / mol } NH_4NO_3} \times 100\% = \frac{4.0316 \text{ g H}}{80.0432 \text{ g } NH_4NO_3} \times 100\%$$

$$= 5.0368\% \text{ H in } NH_4NO_3$$

% O = 100% – 34.9979% C – 5.0368% H = 59.9654% O in NH_4NO_3

☑ *Reasonable Result Check:* Calculating the last element's mass percent using the formula gives the same answer as subtracting the other percentages from 100%.

88. *Result:* **48.203% Fe in $FeCO_3$, 69.9426% Fe in Fe_2O_3, 72.3591% Fe in Fe_3O_4**

Analyze: Given formulas of three compounds, determine the percentage of iron in each of them.

Plan: Calculate the mass of Fe in one mole of compound, while calculating the molar mass of the compound. Divide the calculated element mass by the molar mass of the compound and multiply by 100% to get percent.

Execute: For $FeCO_3$: Mass of Fe per mole of $FeCO_3$ = 55.845 g/mol Fe

Molar mass $FeCO_3$ = (55.845 g/mol Fe) + (12.0107 g/mol C) + 3(15.9994 g/mol O) = 115.856 g/mol $FeCO_3$

$$\%\ \text{Fe} = \frac{\text{mass of Fe/mol FeCO}_3}{\text{mass of FeCO}_3\text{/mol FeCO}_3} \times 100\% = \frac{55.845\ \text{g Fe}}{115.856\ \text{g FeCO}_3} \times 100\% = 48.203\%\ \text{Fe in FeCO}_3$$

For Fe_2O_3: Mass of Fe per mole of Fe_2O_3 = 2(55.845 g/mol Fe) = 111.690 g/mol Fe

Molar mass Fe_2O_3 = 111.690 g/mol Fe + 3(15.9994 g/mol O) = 159.688 g/mol Fe_2O_3

$$\%\ \text{Fe} = \frac{\text{mass of Fe/mol Fe}_2\text{O}_3}{\text{mass of Fe}_2\text{O}_3\text{/mol Fe}_2\text{O}_3} \times 100\% = \frac{111.694\ \text{g Fe}}{159.688\ \text{g Fe}_2\text{O}_3} \times 100\% = 69.9426\%\ \text{Fe in Fe}_2\text{O}_3$$

For Fe_3O_4 Mass of Fe per mole of Fe_3O_4 = 3(55.847 g/mol Fe) = 167.535 g/mol Fe

Molar mass Fe_3O_4 = 167.535 g/mol Fe + 4(15.9994 g/mol O) = 231.533 g/mol

$$\%\ \text{Fe} = \frac{\text{mass of Fe/mol Fe}_3\text{O}_4}{\text{mass of Fe}_3\text{O}_4\text{/mol Fe}_3\text{O}_4} \times 100\% = \frac{167.535\ \text{g Fe}}{231.533\ \text{g Fe}_3\text{O}_4} \times 100\% = 72.3591\%\ \text{Fe in Fe}_3\text{O}_4$$

☑ *Reasonable Result Check:* The percentage of iron increases as the formula includes more iron and less of other elements.

89. *Result:* **245.745 g/mol, 25.858% Cu, 22.7992% N, 5.74197% H, 13.048% S, 32.5528% O**

Analyze: Given the formula of a compound, determine the molar mass, and the mass percent of each element

Plan: Calculate the mass of each element in one mole of compound, while calculating the molar mass of the compound (see full method on Question 68). Divide the calculated mass of the element by the molar mass of the compound and multiply by 100% to get mass percent.

Execute: The compound is $Cu(NH_3)_4SO_4{\cdot}H_2O$.

Mass of Cu per mole of compound = 63.546 g/mol Cu

Mass of N per mole of compound = 4(14.0067 g/mol N) = 56.0268 g/mol N

$Cu(NH_3)_4SO_4{\cdot}H_2O$ contains four NH_3 and one H_2O molecule so it has a total of (3×4 + 2) 14 H atoms.

Mass of H per mole of compound = 14(1.0079 g/mol H)= 14.1106 g/mol H

Mass of S per mole of compound = 32.065 g/mol S

$Cu(NH_3)_4SO_4{\cdot}H_2O$ contains one SO_4 and one H_2O molecule so it has a total of (4 + 1) 5 O atoms.

Mass of O per mole of compound = 5(15.9994 g/mol O) = 79.9970 g/mol O

Molar mass $Cu(NH_3)_4SO_4{\cdot}H_2O$ = (63.546 g/mol Cu) + (56.0268 g/mol N) + (14.1106 g/mol H)

+ (32.065 g/mol S) + (79.9970 g/mol O) = 245.745 g/mol $Cu(NH_3)_4SO_4{\cdot}H_2O$ (comp)

$$\%\ \text{element} = \frac{\text{mass of element/mol comp}}{\text{mass of comp/mol comp}} \times 100\%$$

$$\text{Mass percent Cu} = \frac{63.546\ \text{g Cu}}{245.745\ \text{g comp}} \times 100\% = 25.858\%\ \text{Cu in Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}$$

$$\text{Mass percent N} = \frac{56.0268\ \text{g N}}{245.745\ \text{g comp}} \times 100\% = 22.7992\%\ \text{N in Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}$$

$$\text{Mass percent H} = \frac{14.1106 \text{ g H}}{245.745 \text{ g comp}} \times 100\% = 5.74197\% \text{ H in } Co(NH_3)_4SO_4 \cdot H_2O$$

$$\text{Mass percent S} = \frac{32.065 \text{ g S}}{245.745 \text{ g comp}} \times 100\% = 13.048\% \text{ S in } Co(NH_3)_4SO_4 \cdot H_2O$$

$$\text{Mass percent O} = \frac{79.9970 \text{ g O}}{245.745 \text{ g comp}} \times 100\% = 32.5528\% \text{ O in } Co(NH_3)_4SO_4 \cdot H_2O$$

☑ *Reasonable Result Check:* The sum of the percentages is 100%.

90. *Result:* **291.0678 g/mol, 20.2472% Co, 9.62436% N, 65.96154% O, 4.16686% H, 37.1476% H_2O**

Analyze: Given formulas of a compound, determine the molar mass and mass percent of each element.

Plan: Calculate the mass of each element in one mole of $Co(NO_3)_2{\cdot}6H_2O$, then use those masses to calculate the molar mass of the compound. Divide the calculated mass of each element by the molar mass of the compound and multiply by 100% to get percent.

Execute: The compound is $Co(NO_3)_2{\cdot}6H_2O$.

Mass of Co per mole of compound = 58.9332 g/mol Co

Mass of N per mole of compound = 2(14.0067 g/mol N) = 28.0134 g/mol N

Mass of H per mole of compound = 12(1.0107 g/mol H) = 12.1284 g/mol H

Mass of O per mole of compound = (6 + 6)(15.9994 g/mol O) = 191.9928 g/mol O

Molar mass of $Co(NO_3)_2{\cdot}6H_2O$ = (58.9332 g/mol Co) + (28.0134 g/mol N)

+ (2.1284 g/mol H) + (191.9928 g/mol O) = 291.0678 g/mol

$$\% \text{ element} = \frac{\text{mass of element/mol comp}}{\text{mass of comp/mol comp}} \times 100\%$$

$$\text{Mass percent Co} = \frac{58.9332 \text{ g Co}}{291.0678 \text{ g comp}} \times 100\% = 20.2472\ \% \text{ Co in } Co(NO_3)_2{\cdot}6H_2O$$

$$\text{Mass percent N} = \frac{28.0134 \text{ g N}}{291.0678 \text{ g comp}} \times 100\% = 9.62436\ \% \text{ N in } Co(NO_3)_2{\cdot}6H_2O$$

$$\text{Mass percent O} = \frac{191.9928 \text{ g O}}{291.0678 \text{ g comp}} \times 100\% = 65.96154\ \% \text{ O in } Co(NO_3)_2{\cdot}6H_2O$$

$$\text{Mass percent H} = \frac{12.1284 \text{ g H}}{291.0678 \text{ g comp}} \times 100\% = 4.16686\ \% \text{ H in } Co(NO_3)_2{\cdot}6H_2O$$

$$\text{Mass percent water} = = \frac{6[2(1.0079 \text{ g H}) + (15.9994 \text{ g O})]}{291.0678 \text{ g comp}} \times 100\% = 37.1476\ \%\ H_2O \text{ in } Co(NO_3)_2{\cdot}6H_2O$$

☑ *Reasonable Result Check:* The sum of the mass percents for the elements is 100%.

91. *Result:* **(a) $C_{10}H_{12}NO$ (b) $C_{20}H_{24}N_2O_2$**

Analyze: Given the percent by mass of elements in quinine and the molar mass determine the empirical formula and the molecular formula.

Plan and Execute:

(a) Choose a convenient sample of quinine, such as 100.00 g. Using the mass percents, determine the number of grams of C, H, N, and O in the sample. Convert these masses to moles, using the atomic weights. Set up a ratio, and simplify it by dividing each by the smallest number. Use the integers as the subscripts in the empirical formula

The compound is 74.04% C, 7.46% H, 8.64% N, and 9.86% O by mass. This means that a sample of

100.00 g has 74.04 g C, 7.46 g H, 8.64 g N, and 9.86 g O.

$$74.04 \text{ g C} \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 6.165 \text{ mol C} \qquad 7.46 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 7.40 \text{ mol H}$$

$$8.64 \text{ g N} \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 0.617 \text{ mol N} \qquad 9.86 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.616 \text{ mol O}$$

Mole Ratio 6.16 mol C : 7.40 mol H : 0.617 mol N : 0.616 mol O

Simplify by dividing each amount by 0.616 mol

Mole Ratio 10 C : 12 mol H : 1 mol N : 1 mol O

Therefore, the empirical formula is $C_{10}H_{12}NO$

(b) The molecular formula is $(C_{10}H_{12}NO)_n$. Determine the empirical formula molar mass. Then determine the value of n by dividing the molar mass of the compound by the molar mass of the empirical formula.

Molar mass of $C_{10}H_{12}NO$ = 10(12.0107 g/mol C) + 12(1.0079 g/mol H)

+ 15.9994 g/mol O + 14.0067 g/mol N = 162.2079 g/mol $C_{10}H_{12}NO$

$$n = \frac{324.41 \text{ g/mol comp}}{162.2079 \text{ g/mol emp formula}} = 2$$

So, the molecular formula is $(C_{10}H_{12}NO)_2$, or $C_{20}H_{24}N_2O_2$.

☑ *Reasonable Result Check:* The mole ratio is clearly a whole number relationship and the molar mass is very close to double the empirical formula mass.

92. *Result:* **(a) C_6H_6NOAs (b) $C_{18}H_{18}N_3O_3As_3$**

Analyze: Given the percent by mass of elements in Salvarsan-606 and the molar mass determine the empirical formula and the molecular formula.

Plan and Execute:

(a) Choose a convenient sample of Salvarsan-606, such as 100.00 g. Using the mass percents, determine the number of grams of C, H, N, O, and As in the sample. Convert these masses to moles, using the atomic weights. Set up a ratio, and simplify it by dividing each by the smallest number. Use the integers as the subscripts in the empirical formula

The compound is 39.37% C, 3.304% H, 7.653% N, 8.741% O, and 40.93% As by mass. This means that a sample of 100.00 g has 39.37g C, 3.304 g H, 7.653 g N, 8.741 g O, and 40.93 g As.

$$39.37 \text{ g C} \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 3.278 \text{ mol C} \qquad 3.304 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 3.278 \text{ mol H}$$

$$7.653 \text{ g N} \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 0.5464 \text{ mol N} \qquad 8.741 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.5463 \text{ mol O}$$

$$40.93 \text{ g As} \times \frac{1 \text{ mol As}}{74.9216 \text{ g As}} = 0.5463 \text{ mol As}$$

Mole Ratio 3.278 mol C : 3.278 mol H : 0.5464 mol N : 0.5463 mol O : 0.5463 mol As

Simplify by dividing each amount by : 0.5463 mol

Mole Ratio 6 C : 6 mol H : 1 mol N : 1 mol O : 1 mol As

Therefore, the empirical formula is C_6H_6NOAs

(b) The molecular formula is $(C_6H_6NOAs)_n$. Determine the empirical formula molar mass. Then determine the value of n by dividing the molar mass of the compound by the molar mass of the empirical formula.

Molar mass of C_6H_6NOAs = 6(12.0107 g/mol C) + 6(1.0079 g/mol H)

+ 14.0067 g/mol N + 15.9994 g/mol O + 74.9216 g/mol As = 183.0393 g/mol C_6H_6NOAs

$$n = \frac{549.102\text{ g/mol comp}}{183.0393\text{ g/mol emp formula}} = 3$$

So, the molecular formula is $(C_6H_6NOAs)_3$, or $C_{18}H_{18}N_3O_3As_3$.

☑ *Reasonable Result Check:* The mole ratio is clearly a whole number relationship and the molar mass is very close to triple the empirical formula mass.

93. *Result:* U_3O_8

Analyze: Given the percent by mass of one elements in a uranium oxide, U_xO_y, determine the empirical formula.

Plan and Execute: Choose a convenient sample of U_xO_y, such as 100.00 g. Using the mass percent of U, calculate the mass percent O, then determine the number of grams of U and O in the sample. Convert these masses to moles, using the atomic weights. Set up a ratio, and simplify it by dividing each by the smallest number. Use the integers as the subscripts in the empirical formula

The compound is 84.80% U by mass. The rest is O, 100.00% – 84.80% U = 15.20%. This means that a sample of 100.00 g has 84.80 g U and 15.20 g O.

$$84.80\text{ g U} \times \frac{1\text{ mol U}}{238.0289\text{ g U}} = 0.3563\text{ mol U} \qquad 15.20\text{ g O} \times \frac{1\text{ mol O}}{15.9994\text{ g O}} = 0.9500\text{ mol O}$$

Mole Ratio 0.3563 mol U : 0.9500 mol O

Simplify by dividing each amount by : 0.3563 mol

Mole Ratio 1 U : 2.667 mol O

Multiply by 3, to get whole number:

Therefore, the empirical formula is U_3O_8

☑ *Reasonable Result Check:* The mole ratio is clearly a whole number relationship.

94. *Result:* **One**

Analyze: Given the molar mass of a compound with an unknown formula and the mass percent of an element in that compound, determine the number of atoms of that element in the compound.

Plan Using a convenient sample size of compound, determine the mass of the element in that compound. Convert both masses to moles, then determine the mole ratio.

Execute: In 100.000 grams of the compound carbonic anhydrase (abbreviated as: c.a.) there are 0.218 g Zn.

$$100.000\text{ g c.a.} \times \frac{1\text{ mole c.a.}}{3.00 \times 10^4\text{ g c.a.}} = 0.00333\text{ mole c.a.}$$

$$0.218\text{ g Zn} \times \frac{1\text{ mole Zn}}{65.409\text{ g Zn}} = 0.00333\text{ mole Zn}$$

0.00333 mole Zn : 0.00333 mole c.a.

1 mole Zn : 1 mole c.a.

One Zn atom in every molecule of c.a.

☑ *Reasonable Result Check:* The contribution of the zinc mass to the mass of the molecule is very small, so it makes sense that the number of atoms of zinc is very small in this large molecule.

95. *Result:* **2.20 × 10^5 g/mol**

Analyze: Plan and Execute: Given the percent by mass of an element in an enzyme and the number of atoms of that element in one molecule of the enzyme, determine the molar mass of the enzyme.

Plan: Choose a convenient sample of the enzyme, such as 100.0 g. Using the percent by mass, determine the number of grams of Mo in the sample. Use the molar mass of Mo as a conversion factor to get the moles of Mo. Use the formula stoichiometry as a conversion factor to get the moles of enzyme. Determine the molar mass by dividing the grams of enzyme in the sample, by the moles of enzyme in the sample.

Execute: The enzyme contains 0.0872% Mo by mass, this means that 100.0 g of enzyme contains 0.0872 grams Mo.

Formula stoichiometry: One molecule of enzyme contains 2 atoms of Mo, so 1 mol of enzyme molecules contains 2 mol of Mo atoms.

$$0.0872 \text{ g Mo} \times \frac{1 \text{ mol Mo}}{95.94 \text{ g Mo}} \times \frac{1 \text{ mol enzyme}}{2 \text{ mol Mo}} = 4.54 \times 10^{-4} \text{ mol enzyme}$$

$$\text{Molar Mass of enzyme} = \frac{\text{mass of enzyme in sample}}{\text{moles of enzyme in sample}} = \frac{100.0000 \text{ g enzyme}}{4.54 \times 10^{-4} \text{ mol enzyme}} = 2.20 \times 10^5 \text{ g/mol}$$

☑ *Reasonable Result Check:* Enzymes are large molecules with large molar masses.

96. *Result:* **6**

Analyze: Given the percent by mass of an element in a compound and the compound's formula with an unknown subscript, determine the value of the unknown subscript.

Plan: Choose a convenient sample of Si_2H_x, such as 100.00 g. Using the percent by mass, determine the number of grams of Si and H in the sample. Use the molar mass of Si as a conversion factor to get the moles of Si. Use the molar mass of H as a conversion factor to get the moles of H. Set up a mole ratio to find the value of x.

Execute: The compound is 90.28% Si by mass. This means that 100.00 g of Si_2H_x contains 90.28 grams Si and the rest of the mass is from H.

$$\text{Mass of H in sample} = 100.00 \text{ g } Si_2H_x - 90.28 \text{ g Si} = 9.72 \text{ g H}$$

$$90.28 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.0855 \text{ g Si}} = 3.214 \text{ mol Si}$$

$$9.72 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 9.64 \text{ mol H}$$

$$\text{Mole Ratio} = \frac{\text{moles of H in sample}}{\text{moles of Si in sample}} = \frac{9.64 \text{ mol H}}{3.214 \text{ mol Si}} = \frac{3 \text{ mol H}}{1 \text{ mol Si}} = \frac{6 \text{ mol H}}{2 \text{ mol Si}}$$

Therefore, the formula is Si_2H_6 and x = 6.

☑ *Reasonable Result Check:* The Mole Ratio is clearly a whole number relationship indicating a sensible number of hydrogen atoms in this molecule.

97. *Result:* **(a) 23.6190% (b) No**

(a) *Analyze:* Given the formula of a hydrate compound, and the formula of the hydrate produced after some of the water has been removed, determine the percentage of mass lost during dehydration.

Plan: Find the molar mass of the original hydrate compound. Determine the number of moles of water it lost, and use the molar mass of water to determine mass of water lost per mole of original hydrate compound. Divide the calculated water mass by the molar mass of the compound and multiply by 100% to get percent mass lost.

Execute: The original hydrate compound is $Na_2B_4O_7 \cdot 10\ H_2O$.

Molar Mass of $Na_2B_4O_7 \cdot 10\ H_2O$ = 2(22.9898 g/mol Na) + 4(10.811 g/mol B)

+ (7+ 10)(15.9994 g/mol O) + 20(1.0079 g/mol H) = 381.371 g/mol $Na_2B_4O_7 \cdot 10\ H_2O$

Dehydrating 1 mol $Na_2B_4O_7 \cdot 10\ H_2O$ forms 1 mol $Na_2B_4O_7 \cdot 5\ H_2O$.

Mol of water lost per mole of $Na_2B_4O_7 \cdot 10\ H_2O$ = 5 mol H_2O

Mass of water in 5 mol H_2O = 10(1.0079 g/mol H) + 5(15.9994 g/mol O) = 90.0760 g in 5 mol H_2O

$$\%\ H_2O\ \text{lost} = \frac{90.0760\ \text{g}\ H_2O/\text{mol hydrate}}{381.371\ \text{g hydrate}/\text{mol hydrate}} \times 100\% = 23.6190\%\ H_2O\ \text{lost}$$

(b) *Explanation:* The percent boron by mass will not be the same in these two compounds. Clearly, the number of atoms of boron is the same; however, the numbers of other atoms are significantly different (due to the different amounts of water in the hydrate). The product hydrate, $Na_2B_4O_7 \cdot 5\ H_2O$, will have a larger percent by mass of boron than the original hydrate, $Na_2B_4O_7 \cdot 10\ H_2O$, since there are 5 water molecules fewer in the product than the original.

98. *Result:* $\mathbf{C_4H_8N_2O_2}$

Analyze: Given the empirical formula of a compound and the molar mass, determine the molecular formula.

Plan: Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute: The empirical formula is C_2H_4NO, the molecular formula is $(C_2H_4NO)_n$.

Mass of 1 mol C_2H_4NO = 2(12.0107 g/mol C) + 4(1.0079 g/mol H)

+ 14.0067 g/mol N + 15.9994 g/mol O = 58.0591 g/mol C_2H_4NO

$$n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of } C_2H_4NO} = \frac{116.1\ \text{g}}{58.0591\ \text{g}} = 2.000 \approx 2$$

Molecular Formula is $(C_2H_4NO)_2 = C_4H_8N_2O_2$

$\boxed{\checkmark}$ *Reasonable Result Check:* The molar mass is about 2 times larger than the mass of one mole of the empirical formula, so the molecular formula $C_4H_8N_2O_2$ makes sense.

99. *Result:* $\mathbf{C_3H_4O_3}$

Analyze, Plan, and Execute: An empirical formula shows the simplest whole number ratio of the elements in a compound. The molecular formula gives the actual number of atoms of each element in one formula unit of the compound. For ascorbic acid, $C_6H_8O_6$ is the molecular formula. $\mathbf{C_3H_4O_3}$ is the empirical formula, since both 6 and 8 are exactly divisible by 2 to give the smallest whole number ratio of 3:4:3.

100. *Result:* **x = 12**

Analyze: The mass of a sample of a hydrate compound is given. The formula of the hydrate is known except for the amount of water in it. All the water is dried out of the sample using high temperature, losing a mass of water. Determine the number of moles of water in the formula of the hydrate compound.

Plan: The difference between the mass of the hydrated compound and the mass of the water lost by the sample gives the mass of dehydrated compound. Use the molar mass of the dehydrated compound as a conversion factor to convert the mass of the dehydrated compound into moles. Since the only thing lost was water, a mole relationship can be established between the dehydrated and hydrated compounds. Convert the mass of water into moles, using the molar mass of water. Divide the moles of water by the moles of hydrate, to determine how many moles of water came from one mole of hydrate compound.

Execute: A sample of 4.74 g of hydrated compound, $KAl(SO_4)_2{\cdot}xH_2O$, is dehydrated, losing 2.16 g H_2O, producing $KAl(SO_4)_2$.

Mass of dehydrated compound produced from the sample

$$= 4.74 \text{ g } KAl(SO_4)_2{\cdot}xH_2O - 2.16 \text{ g } H_2O = 2.58 \text{ g } KAl(SO_4)_2$$

Molar Mass $KAl(SO_4)_2$ = 39.0983 g/mol K + 26.9815 g/mol Al + 2(32.065 g/mol S)

$$+ 8(15.9994 \text{ g/mol O}) = 258.206 \text{ g/mol } KAl(SO_4)_2$$

Molar Mass H_2O = 2(1.0079 g/mol H) + 2(15.9994 g/mol O) = 18.0152 g/mol H_2O

Find moles of $KAl(SO_4)_2{\cdot}xH_2O$ in sample:

$$2.58 \text{ g } KAl(SO_4)_2 \times \frac{1 \text{ mol } KAl(SO_4)_2}{258.206 \text{ g } KAl(SO_4)_2} \times \frac{1 \text{ mol } KAl(SO_4)_2 \cdot xH_2O}{1 \text{ mol } KAl(SO_4)_2} = 9.99 \times 10^{-3} \text{ mol } KAl(SO_4)_2{\cdot}xH_2O$$

Find moles of H_2O lost from sample: $2.16 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0152 \text{ g } H_2O} = 0.120 \text{ mol } H_2O$

$$\text{Mole ratio} = \frac{\text{mol } H_2O \text{ from sample}}{\text{mol hydrate is sample}} = \frac{0.120 \text{ mol } H_2O}{9.99 \times 10^{-3} \text{ mol hydrate}} = 12.0 \approx 12$$

The proper formula of the hydrated compound is: $KAl(SO_4)_2{\cdot}12H_2O$ and x = 12.

☑ *Reasonable Result Check:* The mole ratio is very close to a whole number.

General Questions

101. *Result:* **255 mL**

Analyze: A known mass of sulfuric acid is in a solution with a given density and mass percentage. Determine the volume in mL.

Plan: Start with the sample—here, the mass of sulfuric acid. Use the mass percentage as a conversion factor between grams of sulfuric acid and grams of solution. Then use the density of the solution as a conversion factor between grams and cubic centimeters. Then convert the cubic centimeters into milliliters.

Execute: 100 grams of the solution contains 30.08 grams H_2SO_4. 1 cubic centimeter of the solution weighs 1.285 grams.

$$125 \text{ g } H_2SO_4 \times \frac{100 \text{ g solution}}{38.08 \text{ g } H_2SO_4} \times \frac{1 \text{ cm}^3 \text{ solution}}{1.285 \text{ g solution}} \times \frac{1 \text{ mL solution}}{1 \text{ cm}^3 \text{ solution}} = 255 \text{ mL solution}$$

☑ *Reasonable Result Check:* The units "g H_2SO_4" cancel properly, as do the units "g solution". Multiplying a number by 100, then dividing by approximately 50 (= 38 × 1.3), should give an answer about twice as big.

102. *Result:* **0.995 g Pt**

Analyze: Given the mass of a sample of cisplatin along with its percentage platinum, determine the grams of platinum.

Plan: Always start with the sample. Convert the mass of compound to mass of platinum using the percentage platinum as a conversion factor.

Execute: 100 grams of cisplatin contains 65.0 grams of Pt.

$$1.53 \text{ g cisplatin} \times \frac{65.0 \text{ g Pt}}{100 \text{ g cisplatin}} = 0.995 \text{ g Pt}$$

☑ *Reasonable Result Check:* The mass of Pt should be about $\frac{2}{3}$ of the mass of the compound cisplatin.

103. *Result:* **3.24 L**

Analyze: Given the mass and density of C_2H_5OH and the weight of water per liter, determine the volume of C_2H_5OH with the same number of molecules as the water.

Plan and Execute:

Find moles of water in 1.00 kg, using the molar mass of water:

Molar mass H_2O = 2(1.0079 g/mol H) + 15.9994 g/mol O = 18.0152 g H_2O

$$1.00\ \text{kg H}_2\text{O} \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \frac{1\,\text{mol H}_2\text{O}}{18.0152\ \text{g}} = 55.51\ \text{mol H}_2\text{O}$$

A sample with 55.51 mol water has as many molecules as 55.51 mol C_2H_5OH.

Calculate the molar mass then use it and the density to find the volume of C_2H_5OH.

2(12.0107 g/mol C) + 5(1.0079 g/mol H) + 15.9994 g/mol O = 46.068 g C_2H_5OH

$$55.51\ \text{mol C}_2\text{H}_5\text{OH} \times \frac{46.0682\ \text{g C}_2\text{H}_5\text{OH}}{1\ \text{mol C}_2\text{H}_5\text{OH}} \times \frac{1\ \text{mL C}_2\text{H}_5\text{OH}}{0.789\ \text{g C}_2\text{H}_5\text{OH}} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 3.24\ \text{L C}_2\text{H}_5\text{OH}$$

☑ *Reasonable Result Check:* C_2H_5OH has a molar mass greater than water and a density less than water, so it makes sense that the volume of C_2H_5OH is greater than 1 L.

104. *Result:* **7.3 × 10^2 g phosphorus-containing compound, 3.2 × 10^2 g P**

Analyze: Given the mass of a bag of fertilizer along with mass percentage of nitrogen-containing compounds, phosphorus-containing compounds, and potassium-containing compounds in the fertilizer and the mass percentage of phosphorus in the phosphorus-containing compounds, determine the grams of phosphorus-containing compounds.

Plan: Always start with the sample. Convert the bag's mass in pounds to grams. Then use the percentage of phosphorus-containing compounds in the fertilizer as a conversion factor to determine the mass of phosphorus-containing compounds in the bag. Then use the mass percentage of phosphorus in the phosphorus-containing compounds to determine how much phosphorus is in the bag.

Execute: 100 grams fertilizer contains 4.0 grams of phosphorus-containing compounds (PCCs).

$$40.0\ \text{lb fertilizer} \times \frac{453.59\ \text{g fertilizer}}{1\ \text{lb fertilizer}} \times \frac{4.0\ \text{g PCCs}}{100\ \text{g fertilizer}} = 7.3 \times 10^2\ \text{g PCCs}$$

$$7.3 \times 10^2\ \text{g PCCs} \times \frac{43.64\ \text{g P}}{100\ \text{g PCCs}} = 3.2 \times 10^2\ \text{g P}$$

☑ *Reasonable Result Check:* The significant figures are limited to two by the 4.0% figure. The mass units cancel appropriately. The mass of phosphorus is less than half the mass of PCCs.

105. *Result:* **89 tons/yr**

Analyze and Plan: Start with the sample. Given the number of people in the city, use the volume of water each person needs per day, then calculate the total water needs for the day. Using the number of days in a year calculate the total volume of water used per year. Then using the mass of one gallon of water as a conversion factor, determine the grams of water. Convert the grams to tons. Then, using the fluoride concentration as a conversion factor, determine the number of tons of fluoride, and use the mass percentage of fluoride in sodium fluoride to determine the number of tons.

Execute:

$$150{,}000\ \text{people} \times \frac{175\ \text{gal water / person}}{1\ \text{day}} \times \frac{365\ \text{days}}{1\ \text{year}} \times \frac{8.34\ \text{lb water}}{1\ \text{gal water}} \times \frac{1\ \text{ton water}}{2000\ \text{lb water}}$$

$$\times \frac{1\ \text{ton fluoride}}{1{,}000{,}000\ \text{tons water}} \times \frac{100\ \text{tons sodium fluoride}}{45.0\ \text{tons flouride}} = 89\ \frac{\text{tons sodium fluoride}}{\text{year}}$$

☑ *Reasonable Result Check:* The significant figures are limited to two by the 150,000 figure. The mass units are appropriately labeled. The units cancel appropriately to give tons per year. This is a large number of people using a large amount of water so the large quantity of sodium fluoride makes sense.

106. *Result/Explanation:* The symbol ^{37}**Cl** conveys more information than the symbol $_{17}$Cl. All isotopes of chlorine have an atomic number of 17, but only the specific isotope chlorine-37 has a mass number of 37.

107. *Result:* **0.038 mol**

Analyze: Given the carat mass of a diamond and the relationship between carat and milligrams, determine how many moles of carbon are in the diamond.

Plan: Always start with the sample. Diamond is an allotropic form of pure carbon. Given the carats of the diamond, use the relationship between carats and milligrams as a conversion factor to determine milligrams of carbon. Then using metric relationships to determine grams of carbon, and the molar mass of carbon to determine the moles of carbon.

Execute:

$$2.3\text{ carats C}\times\frac{200.\text{ mg C}}{1\text{ carat C}}\times\frac{1\text{ g C}}{1000\text{ mg C}}\times\frac{1\text{ mol C}}{12.01\text{ g C}}=0.038\text{ mol C}$$

☑ *Reasonable Result Check:* A carat is 0.2 grams, so 2.3 carats is less than a half a gram of carbon. Since 12 grams of carbon represents a mole, half a gram should be a few hundredths of a mole.

108. *Result:* **9.08×10^7 g, 1.43×10^6 mol**

Analyze: Given the mass of copper in the Statue of Liberty and the relationship between pounds and grams, determine the total mass (in grams) and amount (in moles) of copper in the Statue of Liberty.

Plan: The sample is the 2.00×10^5 lb copper. Use the given relationship between grams and pounds to determine grams of copper, then use the molar mass of copper to determine the number of moles of copper.

Execute:

$$2.00\times10^{5}\text{ lb Cu}\times\frac{454\text{ g Cu}}{1\text{ lb Cu}}=9.08\times10^{7}\text{ g Cu}$$

$$9.08\times10^{7}\text{ g Cu}\times\frac{1\text{ mol Cu}}{63.55\text{ g Cu}}=1.43\times10^{6}\text{ mol Cu}$$

☑ *Reasonable Result Check:* Gram is a smaller unit of mass than pound, so the mass in grams should be a larger number. The number of moles should be smaller than the number of grams. The number of pounds has three significant figures, so the answer must be reported with three significant figures. The units cancel properly.

109. *Result:* **(a) iodine monobromide (b) bromine trifluoride (c) diiodine hexachloride (d) chlorine pentafluoride (e) iodine heptafluoride**

Analyze and Plan: A general rule for applying the names of binary compounds to the formula is to list the symbol for first element named then the symbol for the second element. Use the prefixes described in Table 2.6 to learn the number of a particular kind of atom and use that number for the subscript on the symbol.

Execute:

(a) IBr has one I atom and one Br atom, so its name is **iodine monobromide**.

(b) BrF_3 has one Br atom and three F atoms, so its name is **bromine trifluoride**.

(c) I_2Cl_6 has two I atom and six Cl atoms, so its name is **diiodine hexachloride**.

(d) ClF_5 has one Cl atom and five F atoms, so its name is **chlorine pentafluoride**.

(e) IF_7 has one I atom and seven F atoms, so its name is **iodine heptafluoride**.

110. *Result:* **(a) See explanation below (b) 4.03298 u (c) 6.69692×10^{-24} g (d) 4.03298 g/mol, which is larger than that on the periodic table, suggesting assumptions are invalid**

Analyze, Plan, and Execute:

(a) The proton and neutron are together in a nucleus and the electrons surround the nucleus.

(b) Table 2.1 gives the masses of the subatomic particles in unified atomic mass units:

Mass of helium = $2(1.00728 \text{ u p}^+) + 2(0.000548579 \text{ u e}^-) + 2(1.00866 \text{ u n}^0) = \mathbf{4.03298\ u}$

(c) Section 2-3 identifies 1 u = 1.66054×10^{-24} g. $4.03298 \text{ u} \times \left(\dfrac{1.66054 \times 10^{-24}\text{ g}}{1 \text{ u}}\right) = 6.69692 \times 10^{-24}$ g

(d) Use Avogadro's number to convert from atoms to moles of atoms.

$$\text{Mass in grams of 1 mole of He atoms} = \left(\frac{6.69692 \times 10^{-24}\text{ g}}{1 \text{ atom}}\right) \times \left(\frac{6.02214179 \times 10^{23}\text{ atoms}}{1 \text{ mol}}\right) = 4.03298 \text{ g/mol}$$

The calculated mass, 4.03298 g/mol, is greater than the mass of 4.0026 g/mol listed on the periodic table, suggesting that it is not appropriate to assume that there is no change in mass of the particles when they are in the atom, as instructed in parts (b) and (c).

111. *Result:* **(a) 28.8515% N (b) 6.57 ×10^{20} molecules (c) 5.26 ×10^{21} C atoms (d) nine times greater**

Analyze: Given the formula caffeine, $C_8H_{10}N_4O_2$, determine the mass percent of nitrogen, molecules of caffeine, and C atoms. Compare caffeine content of beverages.

(a) *Plan:* Calculate the mass of N in one mole of caffeine, while calculating the molar mass of caffeine. Divide the calculated mass of N by the molar mass of caffeine and multiply by 100% to get percent.

Execute: Mass of N per mole of $C_8H_{10}N_4O_2$ = 2(14.0067g N) = 56.0268 g N

Mass of 1 mol $C_8H_{10}N_4O_2$ = 8(12.0107 g/mol C) + 10(1.0079 g/mol H)

+ 4(14.0067 g/mol N) + 2(15.9994 g/mol O) = 194.1902 g/mol $C_8H_{10}N_4O_2$

$$\text{Mass percent N} = \frac{56.0268 \text{ g N}}{194.1902 \text{ g caffeine}} \times 100\% = 28.8515\% \text{ N}$$

(b) *Plan:* Use molar mass and Avogadro's number.

$$\textit{Execute:}\ 212 \text{ mg caffeine} \times \left(\frac{1 \text{ g}}{1000 \text{ mg}}\right) \times \left(\frac{1 \text{ mol caffeine}}{194.1902 \text{ g caffeine}}\right) \times \left(\frac{6.022 \times 10^{23} \text{ caffeine molecules}}{1 \text{ mol caffeine}}\right)$$

$= 6.57 \times 10^{20}$ caffeine molecules

(c) *Plan:* Use the relationship between atoms of C in one molecule.

$$\textit{Execute:}\ 6.57 \times 10^{20} \text{ caffeine molecules} \times \left(\frac{8 \text{ C atoms}}{1 \text{ caffeine molecule}}\right) = 5.26 \times 10^{21} \text{ C atoms}$$

(d) *Plan:* The 1.93-oz 5-Hour Energy® drink has 212 mg caffeine. An 8-oz coffee has 100 mg caffeine. Calculate concentration (mg/oz) and compare.

Execute: Concentration in 5-Hour Energy® drink = $\dfrac{212 \text{ mg}}{1.93 \text{ oz}} = 109.8 \dfrac{\text{mg}}{\text{oz}}$

Concentration in coffee = $\dfrac{100 \text{ mg}}{8 \text{ oz}} = 12.5 \dfrac{\text{mg}}{\text{oz}}$ (1 sig fig)

Compare the concentrations, using a ratio: $\dfrac{109.8 \frac{\text{mg}}{\text{oz}}}{12.5 \frac{\text{mg}}{\text{oz}}} = 8.78 = 9$

The 5-Hour Energy® drink has **nine times greater** caffeine concentration.

☑ *Reasonable Result Check:* This explains why people use energy drinks instead of coffee to keep awake.

112. *Result:* **(a) 4.652×10^{-23} g (b) 5.314×10^{-23} g (c) 1.14227, they are the same**

Analyze: Given one molecule of nitrogen, determine its mass in grams. Given one molecule of oxygen, determine its mass in grams. Find the ratio of the masses of these two atoms and compare that ratio to the ratio of atomic weights of nitrogen and oxygen.

Plan: Use the periodic table to get the atomic weights of nitrogen and oxygen and equate those to the mass in grams of one mole of nitrogen molecules. Divide those numbers by Avogadro's number to get the masses of one nitrogen and one oxygen molecule in grams. Take the ratio and compare to the direct ratio.

Execute:

(a) $$\frac{28.0134 \text{ g } N_2}{1 \text{ mol } N_2 \text{ molecules}} \times \frac{1 \text{ mol } N_2 \text{ molecules}}{6.022 \times 10^{23} \; N_2 \text{ molecules}} = 4.652 \times 10^{-23} \; \frac{\text{g}}{N_2 \text{ molecule}}$$

(b) $$\frac{31.9988 \text{ g } O_2}{1 \text{ mol } O_2 \text{ molecules}} \times \frac{1 \text{ mol } O_2 \text{ molecules}}{6.022 \times 10^{23} \; O_2 \text{ molecules}} = 5.314 \times 10^{-23} \; \frac{\text{g}}{O_2 \text{ molecule}}$$

(c) $$\text{molecule mass ratio} = \frac{5.314 \times 10^{-23} \text{ g } O_2}{4.652 \times 10^{-23} \text{ g } N_2} = 1.14227$$

$$\text{atomic weight ratio} = \frac{31.9994 \text{ amu } O_2}{28.0134 \text{ amu } N_2} = 1.14227$$ The ratios are identical.

☑ *Reasonable Result Check:* The molecules are very small, so their mass should be very small. Avogadro's number is a physical constant, so the ratio of the masses must be the same.

113. *Result:* **(a) Ionic: (ii), (iv), and (vi) Molecular: (i), (v), (vii) No compound forms: (iii) (b) (i) BrCl, bromine monochloride; (ii) Li_2Te, lithium telluride; (iv) MgF_2, magnesium fluoride, (v) NF_3, nitrogen trifluoride; (vi) In_2S_3, indium sulfide; (vii) $SeBr_2$, selenium dibromide**

Analyze and Plan: Identify ionic, molecular or no compound. If a compound forms, determine the compound's formula and name.

Execute: (a) Ionic compound, molecular compound, or no compound. (b) Formula and name.

(i) Chlorine (Cl) and bromine (Br) are not likely to form an ionic compound, since they are both nonmetals in Group 7A. If a **molecular** compound formed, it would be covalent **BrCl, bromine monochloride**.

(ii) Lithium (Li) and tellurium (Te) might make an **ionic** compound. Lithium is a metal and tellurium is a metalloid. The likely compound contains ions Li^+ (Group 1A cation; charge is +1) and Te^{2-} (Group 6A anion; charge is –2). The compound's formula will be **Li_2Te, lithium telluride**.

(iii) Sodium (Na) and argon (Ar) are not likely to form an ionic compound, since argon is in Group 8A. Those elements a very unreactive and do not form ions at all. **No compound is expected to form**.

(iv) Magnesium (Mg) and fluorine (F) will make an **ionic** compound. Magnesium is a metal and fluorine is a nonmetal. The likely compound contains ions Mg^{2+} (Group 2A cation; charge is +2) and F^- (Group 7A anion; charge is –1). The compound's formula will be **MgF_2, magnesium fluoride**.

(v) Nitrogen (N) and bromine (Br) are not likely to form an ionic compound, since they are both nonmetals in Groups 5A and 7A, respectively. If a **molecular** compound formed, it would likely be covalent **NF_3, nitrogen trifluoride**.

(vi) Indium (In) and sulfur (S) will make an **ionic** compound. Indium is a metal and sulfur is a nonmetal. The likely compound contains ions In^{3+} (Group 3A cation; charge is +3) and S^{2-} (Group 6A anion; charge is –2). The compound's formula will be **In_2S_3, indium sulfide**.

(vii) Selenium (Se) and bromine (Br) are not likely to form an ionic compound, since they are both nonmetals in Groups 6A and 7A, respectively. If a **molecular** compound formed it would be covalent **$SeBr_2$, selenium dibromide**.

114. *Result:* **(a) (i) NaClO (ii) P_4O_{10} (iii) $KMnO_4$ (iv) KH_2PO_4 (v) ClF_3 (vi) BBr_3 (vii) $Ca(CH_3COO)_2$ (viii) Na_2SO_3 (b) Ionic: (i), (iii), (iv), (vii), and (viii); Molecular: (ii), (v), and (vi)**

Analyze and Plan: Identify ionic or molecular, then determine the compound's formula.

Execute:

(a) Find the type of compound to determine the formulas:

(i) sodium hypochlorite (common cation, Na^+, and anion, ClO^-; **ionic**): **NaClO**

(ii) tetraphosphorus decaoxide (binary molecular compound since both elements are nonmetals; **molecular**). $\mathbf{P_4O_{10}}$

(iii) potassium permanganate (common cation, K^+, and anion, MnO_4^-; **ionic**): $\mathbf{KMnO_4}$

(iv) potassium dihydrogen phosphate (common cation, K^+, and anion, $H_2PO_4^-$; **ionic**): $\mathbf{KH_2PO_4}$

(v) chlorine trifluoride (binary molecular compound since both elements are nonmetals; **molecular**). $\mathbf{ClF_3}$

(vi) boron tribromide (binary molecular compound since both elements are nonmetals; **molecular**). $\mathbf{BBr_3}$

(vii) calcium acetate (common cation, Ca^{2+}, and anion, CH_3COO^- ; **ionic**): $\mathbf{Ca(CH_3COO)_2}$

(viii) sodium sulfite (common cation, Na^+, and anion, SO_3^{2-}; **ionic**): $\mathbf{Na_2SO_3}$

(b) As described in the solution to (a), above, the following are ionic compounds are: (i), (iii), (iv), (vii), and (viii) and the following are molecular compounds are: (ii), (v), and (vi)

115. *Result:* **(a) 1.66×10^{-3} mol (b) 0.346 g**

Analyze: Given the number of tablets consumed, the mass of a compound in each tablet, and the formula of the compound, determine the moles of the compound consumed and the mass of one element consumed.

Plan: Always start with the sample—in this case, the number of tablets consumed. We assume that $C_7H_5BiO_4$ is the "active ingredient". Use the mass of $C_7H_5BiO_4$ per tablet to determine the mass of $C_7H_5BiO_4$ in the sample. Then use the molar mass of $C_7H_5BiO_4$ to determine the moles of $C_7H_5BiO_4$ in the sample. Then use the formula stoichiometry to get the moles of Bi in the sample. Then use the molar mass of Bi to get the grams of Bi in the sample.

Execute: The sample is composed of two tablets of Pepto-Bismol. Find grams of $C_7H_5BiO_4$:

$$2\ \text{tablets} \times \frac{300.\ \text{mg}\ C_7H_5BiO_4}{1\ \text{tablet}} \times \frac{1\ \text{g}\ C_7H_5BiO_4}{1000\ \text{mg}\ C_7H_5BiO_4} = 0.600\ \text{g}\ C_7H_5BiO_4$$

Molar mass of $C_7H_5BiO_4$= 7(12.0107 g/mol C) + 5(1.0079 g/mol H)

+ 208.9804 g/mol Bi + 4(15.9994 g/mol O) = 362.0924 g/mol $C_7H_5BiO_4$

(a) Find moles of $C_7H_5BiO_4$ in the sample

$$0.600\ \text{g}\ C_7H_5BiO_4 \times \frac{1\ \text{mol}\ C_7H_5BiO_4}{362.0924\ \text{g}\ C_7H_5BiO_4} = 1.66 \times 10^{-3}\ \text{mol}\ C_7H_5BiO_4$$

(b) Find grams of Bi in the sample

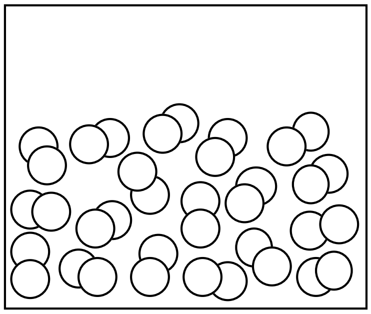
$$1.66 \times 10^{-3}\ \text{mol}\ C_7H_5BiO_4 \times \frac{1\ \text{mol Bi}}{1\ \text{mol}\ C_7H_5BiO_4} \times \frac{208.9804\ \text{g Bi}}{1\ \text{mol Bi}} = 0.346\ \text{g Bi}$$

☑ *Reasonable Result Check:* (a) The sample is somewhere between macroscale and microscale, so it makes sense that the number of moles is somewhat small. (b) The bismuth is almost 60% of the $C_7H_5BiO_4$ compound mass so it makes sense that the number of grams of Bi in the sample is about 60% of the mass of the compound in the sample.

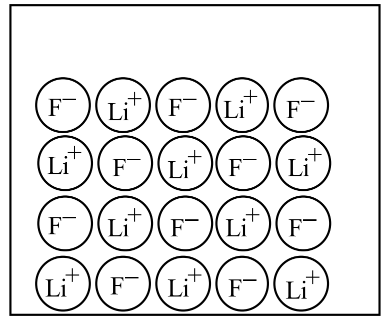
Applying Concepts

116. *Analyze and Plan, and Execute:*

Liquid bromine: Br_2 (ℓ)

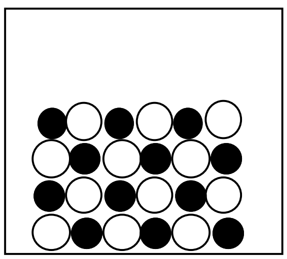


Solid lithium fluoride: LiF (s)

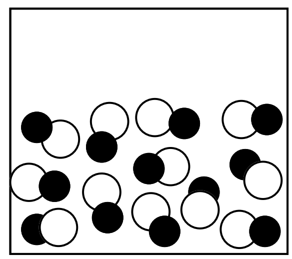


117. *Analyze and Plan, and Execute:*

(a) A crystal of sodium chloride has alternating lattice of Na^+ and Cl^- ions.



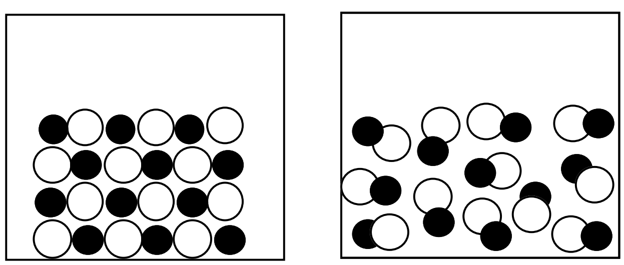
(b) The sodium chloride after it is melted has paired ions randomly distributed.



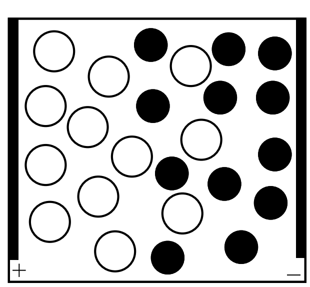
118. *Analyze and Plan, and Execute:*

(a) Solid lithium nitrate has alternating lattice of Li^+ and NO_3^- ions.

(b) Molten lithium nitrate has ion pairs randomly distributed.



(c) Molten lithium nitrate when positive and negative electrodes are present will have the anions, NO_3^-, (white circles) crowding around the positive electrode and the cations, Li^+, (the black circles) crowding around the negative electrode:



119. *Result:* **(a) not possible; mass number and atomic number wrong (b) possible (c) not possible; mass number wrong (d) not possible; mass number wrong (e) possible (f) not possible; mass number wrong**

Analyze and Plan: Check to see if the following statements are true: The atomic number must be the same as the number of protons and the mass number must be the sum of the number of protons and neutrons.

Execute:

(a) These values are **not possible**, since the atomic number (42) is not the same as the number of protons (19), and the sum of protons and neutrons (19+23=42) is not the same as the mass number (19).

(b) These values are **possible**. The atomic number (92) is the same as the number of protons (92), and the sum of protons and neutrons (92+143=235) is the same as the mass number (235).

(c) These values are **not possible**. The sum of protons and neutrons (131+79=210) is not the same as the mass number (53).

(d) These values are **not possible**. The sum of protons and neutrons (15+15=30) is not the same as the mass number (32).

(e) These values are **possible**. The atomic number (7) is the same as the number of protons (7), and the sum of protons and neutrons (7+7=14) is not the same as the mass number (14).

(f) These values are **not possible**. The sum of protons and neutrons (18+40=58) is not the same as the mass number (40).

120. *Result:* **(a) not possible; mass number wrong (b) not possible; atomic number wrong (c) not possible; mass number wrong (d) possible (e) not possible; atomic number wrong**

Analyze and Plan: Check to see if the following statements are true: The atomic number must be the same as the number of protons and the mass number must be the sum of the number of protons and neutrons.

Execute:

(a) These values are **not possible**, since the sum of protons and neutrons (25+29=54) is not the same as the mass number (53).

(b) These values are **not possible**, since the atomic number (78) is not the same as the number of protons (195), and the sum of protons and neutrons (195+117=312) is not the same as the mass number (195).

(c) These values are **not possible**. The sum of protons and neutrons (16+16=32) is not the same as the mass number (33).

(d) These values are **possible**. The atomic number (24) is the same as the number of protons (24), and the sum of protons and neutrons (24+28=52) is not the same as the mass number (52).

(e) These values are **not possible**, since the atomic number (17) is not the same as the number of protons (18).

121. *Result:* $\mathbf{^{39}K}$

Analyze, Plan, Execute: Potassium's atomic weight is 39.0983. The isotopes that contribute most to this mass are ^{39}K and ^{41}K, since the question tells us that ^{40}K has a very low abundance. Since the atomic mass is closer to 39 than 41, that confirms that the $\mathbf{^{39}K}$ **isotope** is more abundant.

122. *Result:* $\mathbf{^{7}Li}$

Analyze, Plan, Execute: The two isotopes of lithium are ^{6}Li and ^{7}Li. The mass of ^{6}Li is close to 6 u and the mass of ^{7}Li is close to 7 u. Because lithium's atomic weight (6.941 u) is much closer to 7 u than to 6 u, the isotopic $\mathbf{^{7}Li}$ **is more abundant** than the isotope ^{6}Li.

123. *Result:* **(a) 1 mol of $\mathbf{Cl_2}$ (b) 1 mol of $\mathbf{O_2}$ (c) one nitrogen molecule (d) $\mathbf{6.032\times10^{23}}$ molecules of $\mathbf{F_2}$ (e) 20.3 grams of neon (f) 159.8 grams $\mathbf{Br_2}$ (g) 9.6 grams of Li (h) 58.9 g Co (i) $\mathbf{6.022\times10^{23}}$ calcium atoms (j) Same**

Analyze and Plan: Molecules are made up of atoms. The unit mole is a convenient way of describing a large quantity of particles. It is also important to keep in mind that 1 mol of particles contains 6.022×10^{23} particles and the molar mass describes the mass of 1 mol of particles.

Execute:

(a) A sample containing **1 mol of Cl_2** has more atoms than a sample containing 1 mol Cl, since each molecule of Cl_2 contains two atoms of Cl.

(b) A sample containing **1 mol of O_2** contains 6.022×10^{23} molecules of O_2. This sample has many more atoms than a sample containing just 1 molecule of O_2.

(c) A sample containing **one nitrogen molecule** (N_2) has two atoms, which is more than one N atom.

(d) A sample containing **6.032×10^{23} molecules of F_2** contains more than 1 mol of F_2 molecules, which has only 6.022×10^{23} molecules of F_2.

(e) The molar mass of Ne is 20.18 g/mol, so a sample of **20.3 grams of neon** contains more than 1 mol of neon.

(f) The molar mass of bromine (Br_2) is 2(79.9 g/mol Br) = 159.8 g/mol Br_2, so the sample composed of **159.8 grams of bromine** contains 1 mol bromine, which equals 6.022×10^{23} molecules. This 159.8-gram sample has many more particles than a sample containing just 1 molecule of Br_2.

(g) The molar mass of Ag is 107.9 g/mol, so a sample of 107.9 grams of Ag contains 1 mol of Ag. The molar mass of Li is 6.9 g/mol, so a sample of **9.6 grams of Li** contains more than 1 mol of Li.

(h) The molar mass of Co is 58.9 g/mol, so the sample with **58.9 grams of Co** contains 1 mol of Co atoms. The molar mass of Cu is 63.55 g/mol, so the sample with 58.9 grams of Cu contains less than 1 mol of Cu atoms. Therefore the Co sample has more atoms than the Cu sample.

(i) The sample containing 6.022×10^{23} atoms of calcium, Ca, contains 1 mol of Ca atoms. The molar mass of Ca is 40.1 g/mol, so the sample composed of 1 gram of cobalt, Co, contains less than 1 mol of Co; thus the sample with **6.022×10^{23} calcium atoms** has more particles than a sample with 1 gram cobalt.

(j) Since chlorine atoms all weigh the same, so two samples with identical mass containing only chlorine must have the **same** number of Cl atoms. The molar mass of Cl_2 is twice that of Cl, but the number of molecule is half the number of atoms, if the samples are both 1 g.

124. *Result:* **(a) 1 mol of Fe (b) 6.022×10^{24} lead atoms (c) 1 mol of copper (d) 1 mol of Cl_2 (e) Same (f) 1 mol of Mg (g) 1 mol of Na (h) 4.1 g of He (i) 4.1 g of He (j) 1 oxygen molecule**

Analyze and Plan: The unit mole is a convenient way of describing a large quantity of particles. It is also important to keep in mind that 1 mol of particles contains 6.022×10^{23} particles and that the molar mass gives the grams in one mol of particles.

Execute:

(a) The molar mass of iron (Fe) is 55.845 g/mol. The molar mass of aluminum (Al) is 27.0 g/mol. So a sample of **1 mol of Fe** has a greater mass than one with 1 mol sample of Al.

(b) A sample of **6.022×10^{24} lead atoms** contains more than 1 mol of lead, which is 6.022×10^{23} lead atoms. This sample will have a greater mass than a sample containing 1 mol of lead.

(c) **1 mol of copper** contains 6.022×10^{23} copper atoms. This sample will have a greater mass than a sample containing only 1 copper atom.

(d) A Cl_2 molecule has twice the mass of one Cl atom. So, comparing samples with the same quantity of particles, **1 mol of Cl_2** sample will have a greater mass than the Cl sample.

(e) A gram is a unit of mass. If both samples weigh 1 gram, then they have the **same** mass.

(f) The molar mass of magnesium (Mg) is 24.3 g/mol, so a sample weighing 23.4 grams contains less than 1 mol of Mg. The sample containing **1 mol of Mg** will have a greater mass than 23.4 g of Mg.

(g) The molar mass of Na is 23.0 g/mol, so a **1 mol sample of Na** will weigh 23.0 grams. This sample has a greater mass than a sample containing only 1 gram of Na.

(h) The molar mass of He is 4.0 g/mol, so a **4.1 g sample of He** weighs more than 1 mole. A sample of 6.022×10^{23} He atoms contains 1 mol of He. These two samples have the same mass.

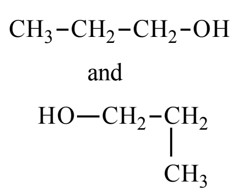
(i) A sample with **1 mol of I_2** contains 6.022×10^{23} I_2 molecules. This sample weight is greater than a sample that only contains 1 I_2 molecule.

(j) **One oxygen molecule** (O_2) has twice the mass of one O atom, so the O_2 sample will have a greater mass than the O sample.

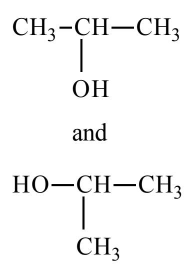
126. *Result:* **(a) three (b) (i) and (iv), (ii) and (iii), (v) and (vi)**

Analyze, Plan, and Execute: There are three isomers given for C_3H_8O. Each of these isomers is represented by two of the structures. The following pairs are identical:

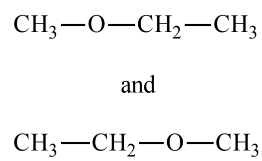
Isomer number one:



Isomer number two:



Isomer number three:



The first pair has an OH bonded to an end carbon: **(i) and (iv)**. The second pair has an OH bonded to the middle carbon: **(v) and (vi)**. The third pair has an O bonded between two carbon atoms: **(ii) and (iii)**.

127. *Result:* **Tl_2CO_3, Tl_2SO_4**

Analyze, Plan, and Execute: Thallium nitrate is $TlNO_3$. Since NO_3^- has a –1 charge, that means that thallium ion has a +1 charge, and is represented by Tl^+. The carbonate compound containing thallium will be a combination of Tl^+ and CO_3^{2-}, and the compound's formula will look like this: Tl_2CO_3. The sulfate compound containing thallium will be a combination of Tl^+ and SO_4^{2-}, and the compound's formula will look like this: Tl_2SO_4.

128. *Result:* **(a) calcium fluoride (b) copper(II) oxide (c) sodium nitrate (d) nitrogen triiodide (e) iron(III) chloride (f) lithium sulfate**

Analyze, Plan, and Execute:

(a) CaF_2 is calcium fluoride. Don't use the "di-" prefix when naming ionic compounds.

(b) CuO is copper(II) oxide. The transition elements have several ions with different charges, so the specific valence is described by a Roman numeral indicating the cation's charge. Since oxide ion has the formula O^{2-}, the copper ion present is, Cu^{2+}, the copper(II) ion.

(c) $NaNO_3$ is sodium nitrate. The incorrect name inappropriately used the naming system for binary molecular compounds, but this compound has more than two elements. It must be named using the ionic compound naming system by naming the common cation (Na^+, sodium) and anion (NO_3^-, nitrate).

(d) NI_3 is a binary molecular compound, containing only two elements. The name is nitrogen triiodide.

(e) $FeCl_3$ is iron(III) chloride. The Roman numeral indicates the charge of the iron cation. Since chloride ion is Cl^-, and three of them are in this neutral compound, that means the iron ion present is Fe^{3+}, which is called iron(III) ion, not iron(I).

(f) Li_2SO_4 is lithium sulfate. The incorrect name inappropriately used the naming system for binary molecular compounds, and this compound has more than two elements. It must be named using the ionic compound naming system by naming the common cation (Li^+, lithium) and anion (SO_4^{2-}, sulfate).

More Challenging Questions

129. *Result:* **5.014×10^{19} atoms, 5.52×10^{19} atoms, 4.5×10^{19} atoms**

NOTE: If you assign this question, explain to the class what "to the nearest 1.0×10^{-4} g" means. It could be interpreted as the uncertainty of the mass measurement is ± 1.0×10^{-4} g (as described in the answer below), or it could be interpreted as the span of uncertainty, so the uncertainty is ± 5×10^{-5} g.

Analyze: Given the mass of a sample of carbon, determine the number of carbon atoms. Given the precision of an instrument that measures mass, determine the high and low limits on the number of atoms present in the sample.

Plan: Convert from mass to moles using the molar mass of carbon, then use Avogadro's number to determine the number of carbon atoms. Based on the uncertainty of the measurement, determine the highest mass and lowest mass that this sample might have. Then calculate the number of carbon atoms in each sample.

Execute:

$$1.000\text{ mg C}\times\frac{1\text{ g}}{1000\text{ mg}}=1.000\times10^{-3}\text{ g C}$$

$$1.000\times10^{-3}\text{ g C}\times\frac{1\text{ mol C}}{12.0107\text{ g C}}\times\frac{6.022\times10^{23}\text{ C atoms}}{1\text{ mol C}}=5.014\times10^{19}\text{ C atoms}$$

The question indicates that the uncertainty of the mass measurement is $\pm1.0 \times 10^{-4}$ g, so the mass may be as large as 1.000×10^{-3} g + 1.0×10^{-4} g = 1.000×10^{-3} g + 0.10×10^{-3} g = 1.10×10^{-3} g and as small as 1.000×10^{-3} g – 1.0×10^{-4} g = 1.000×10^{-3} g – 0.10×10^{-3} g = 0.90×10^{-3} g.

The most atoms that may be present (i.e., the high limit) are:

$$1.10\times10^{-3}\text{ g C}\times\frac{1\text{ mol C}}{12.0107\text{ g C}}\times\frac{6.022\times10^{23}\text{ C atoms}}{1\text{ mol C}}=5.52\times10^{19}\text{ C atoms}$$

The least atoms that may be present (i.e., the low limit) are:

$$0.90\times10^{-3}\text{ g C}\times\frac{1\text{ mol C}}{12.0107\text{ g C}}\times\frac{6.022\times10^{23}\text{ C atoms}}{1\text{ mol C}}=4.5\times10^{19}\text{ C atoms}$$

☑ *Reasonable Result Check:* The variability of the number of atoms is in the last significant figure of the value with the least significant figures, as expected.

130. *Result:* **(a) ^{79}Br–^{79}Br, ^{79}Br–^{81}Br, ^{81}Br–^{81}Br (b) 78.918 g/mol, 80.916 g/mol (c) 79.92 g/mol (d) 50.1% ^{79}Br, 49.9% ^{81}Br**

Analyze: Using the mass of several diatomic molecules of bromine that differ by iostopic composition and the relative heights of their spectral peaks on a mass spectrum, determine the identity of isotopes in the molecules, the mass of each isotope of bromine, the average atomic mass of bromine, and the abundance of the isotopes.

Plan: Identify which of the two isotopes contribute to each of the mass spectrum peaks. Then use that information to find that atomic mass of each isotope. Use the relative peak heights and the isotope masses to find the average molar mass of Br_2 and the average atomic mass of Br, then establish variables describing the isotope percentages. Set up two relationships between these variables. The sum of the percents must be 100%, and the weighted average of the molecular masses must be the reported atomic mass.

Execute:

(a) The peak representing the diatomic molecule with the lightest mass must be composed of two atoms of the lightest isotopes, ^{79}Br–^{79}Br. The peak representing the diatomic molecule with the largest mass must be composed of two atoms of the heavier isotopes, ^{81}Br–^{81}Br. The middle peak must represent molecules made with one of each, ^{79}Br–^{81}Br.

(b) The mass of the molecule composed of two lightweight atoms is 157.836 g/mol, so each atom must weigh 0.5×(157.836 g/mol) = **78.918 g/mol**. The mass of the molecule composed of two heavy weight atoms is 161.832 g/mol, so each atom must weigh 0.5×(161.832 g/mol) = **80.916 g/mol**. This is verified by adding these two isotope masses to obtain the mass of the mixed-isotope peak: 78.918 g/mol + 80.916 g/mol = 159.834 g/mol

(c) The peak heights relate to the quantity of each isotope variation, so find the percentage abundance for each molecule by calculating the percentage of peak heights:

Total = 6.337 + 12.499 + 6.164 = 25.000

$$\%\ {}^{79}\text{Br–}{}^{79}\text{Br} = \frac{6.337}{25.000} \times 100\ \% = 25.35\ \%$$

$$\%\ {}^{79}\text{Br–}{}^{81}\text{Br} = \frac{12.499}{25.000} \times 100\ \% = 50.000\ \%$$

$$\%\ {}^{81}\text{Br–}{}^{81}\text{Br} = \frac{6.164}{25.000} \times 100\ \% = 24.66\ \%$$

The average mass of the molecules would be the weighted average of these three isotopic variants:

$$\frac{25.35 \text{ atoms } {}^{79}\text{Br}_2}{100 \text{ Br}_2 \text{ molecules}} \times \left(\frac{157.836 \text{ u}}{1 \text{ atom } {}^{79}\text{Br}_2}\right) + \frac{50.00 \text{ atoms } {}^{79}\text{Br}{}^{81}\text{Br}}{100 \text{ Br}_2 \text{ molecules}} \times \left(\frac{159.834 \text{ u}}{1 \text{ atom } {}^{79}\text{Br}{}^{81}\text{Br}}\right) +$$

$$\frac{24.66 \text{ atoms } {}^{81}\text{Br}_2}{100 \text{ Br}_2 \text{ molecules}} \times \left(\frac{161.832 \text{ u}}{1 \text{ atom } {}^{81}\text{Br}_2}\right) = 159.84 \frac{\text{u}}{\text{Br}_2 \text{ molecules}}$$

The diatomic molecule mass gets divided by two, to determine the atomic mass:
$(159.84 \text{ g/mol } Br_2)/2 = 79.92$ g/mol Br.

(d) Define percent abundance in terms of two variables: X% ^{79}Br and Y% ^{81}Br. This means:

Every 100 atoms of bromine contains X atoms of the ^{79}Br isotope.

Every 100 atoms of bromine contains Y atoms of the ^{81}Br isotope.

$$X + Y = 100\%$$

$$\frac{\text{X atoms } {}^{79}\text{Br}}{100 \text{ Br atoms}} \times \left(\frac{78.918 \text{ u}}{1 \text{ atom } {}^{79}\text{Br}}\right) + \frac{\text{Y atoms } {}^{91}\text{Br}}{100 \text{ Br atoms}} \times \left(\frac{80.916 \text{ u}}{1 \text{ atom } {}^{91}\text{Br}}\right) = 79.92 \frac{\text{u}}{\text{Br atom}}$$

We now have two equations and two unknowns, so we can solve for X and Y algebraically. Solve the first equation for Y: Y = 100 – X. Plug that in for Y in the second equation. Then solve for X:

$$\frac{X}{100} \times (78.918) + \frac{100 - X}{100} \times (80.916) = 79.92$$

$$0.78918\ X + 80.916 - 0.80916\ X = 79.92$$

$$80.916 - 79.92 = 0.80916\ X - 0.78918\ X$$

$$1.00 = (0.01998)X$$

$$X = 50.1$$

Now, plug the value of X in the first equation to get Y. Y = 100 – X = 100 – 50.1 = 49.9

Therefore the abundance for these isotopes are: **50.1% ^{79}Br and 49.9% ^{81}Br**.

☑ *Reasonable Result Check:* The molar mass of Br from the periodic table is 79.904 g/mol, so the result from part (c) is close but slightly high. These isotopes are about equally abundant as indicated by the near 25%:50%:25% ratio of the isotopes in the three mass spectrum peaks. The % isotopic abundance values calculated in (d) are similar to but not the same as the percentages given in Question 16, 50.69% and 49.31%.

131. *Result:* **(a) 7.10×10^{-4} mol U, U_2O_5, uranium (V) oxide, 3.5×10^{-4} mol U_2O_5 (b) five**

(a) *Analyze:* Given the mass of a uranium oxide sample and the mass of uranium metal in the sample, determine the moles of uranium, the formula of the uranium oxide compound, and the moles of uranium oxide produced.

Plan: The sample is the mass of U_xO_y. Subtract the mass of U from the mass of U_xO_y to find the moles of O in the sample. Then use the molar masses of U and O to find the moles of U and moles of O in the sample. Then set up a mole ratio and find the simplest whole number relationship between the atoms.

Execute: Mass of O in the sample = U_xO_y compound mass – U mass

$$0.199 \text{ g } U_xO_y - 0.169 \text{ g U} = 0.030 \text{ g O}$$

We must keep only three decimal places, resulting in a reduction in the number of sig. fig's in the oxygen mass.

Find moles of U and O in sample:

$$0.169 \text{ g U} \times \frac{1 \text{ mol U}}{238.03 \text{ g U}} = 7.10 \times 10^{-4} \text{ mol U}$$

$$0.030 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 1.9 \times 10^{-3} \text{ mol O}$$

Set up mole ratio:

$$7.10 \times 10^{-4} \text{ mol U} : 1.9 \times 10^{-3} \text{ mol O}$$

Divide all the numbers in the ratio by the smallest number of moles, 7.10×10^{-4} mol: 1 U : 2.6 O

2.6 is close to both 2.66 and 2.5, and since this number has an uncertainty of ± 0.1, either one is equally valid. It is unclear whether we should round down or up to find the whole number ratio. So, look at each case.

If the real ratio is $2\frac{2}{3}$, we multiply by 3, and get 3 U : 8 O and an empirical formula of U_3O_8.

If the real ratio is $2\frac{1}{2}$, we multiply by 2, and get 2 U : 5 O and an empirical formula of U_2O_5.

Of the two empirical formulas of U_3O_8 and U_2O_5, the second formula makes more sense. The common ion of oxygen is oxide ion, O^{2-}. U_2O_5 looks like the simple combination of U^{5+} and O^{2-}. The U_3O_8 formula is either wrong, or contains some mixture of uranium oxides with different formulas, such as a mixture of UO_3 and U_2O_5. With this insight, let's presume that the U_2O_5 formula is the right formula. U_2O_5 would be called uranium (V) oxide.

Molar mass of U_2O_5 = 2(238.03 g/mol U) + 5(15.9994 g/mol O) = 556.06 g/mol U_2O_5

$$0.199 \text{ g } U_2O_5 \times \frac{1 \text{ mol } U_2O_5}{556.06 \text{ g } U_2O_5} = 3.58 \times 10^{-4} \text{ mol } U_2O_5$$

(b) *Analyze:* Given the mass of a sample of a uranium oxide hydrate and the mass of uranium oxide after it is dehydrated, determine the number of molecules of water associated with the hydrate.

Plan: Subtract the mass of the dehydrated compound, $UO_2(NO_3)$, from the mass of the hydrate, $UO_2(NO_3){\cdot}nH_2O$, to get the mass of water. Use the molar mass of water to get the moles of water. Use the molar mass of the $UO_2(NO_3)$ and its relationship to the hydrated compound to find the number of moles of hydrate. Then set up a mole ratio between moles of water and moles of hydrate to find out how many molecules of water are associated with one hydrate formula.

Execute:

Mass of H_2O in the sample = 0.865 g $UO_2(NO_3){\cdot}nH_2O$ – 0.679 g $UO_2(NO_3)$ = 0.186 g H_2O

Molar mass of H_2O = 2(1.0079 g/mol H) + 15.9994 g/mol O = 18.0152 g/mol H_2O

Molar mass of $UO_2(NO_3)$ = 238.03 g/mol U + 5(15.9994 g/mol O)

+ 14.0067 g/mol N = 332.03 g/mol $UO_2(NO_3)$

Mol of water in sample: $0.186 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0152 \text{ g } H_2O} = 0.0103 \text{ mol } H_2O$

Moles of $UO_2(NO_3)\cdot nH_2O$ in sample:

$$0.679 \text{ g } UO_2(NO_3) \times \frac{1 \text{ mol } UO_2(NO_3)}{332.03 \text{ g } UO_2(NO_3)} \times \frac{1 \text{ mol } UO_2(NO_3)\cdot nH_2O}{1 \text{ mol } UO_2(NO_3)} = 0.00204 \text{ mol } UO_2(NO_3)\cdot nH_2O$$

Mole ratio:

$$0.0103 \text{ mol } H_2O : 0.00204 \text{ mol } UO_2(NO_3)\cdot nH_2O$$

Divide by the smallest number of moles, 0.00204 mol, and round to whole numbers:

$$5\ H_2O : 1\ UO_2(NO_3)\cdot nH_2O$$

That means n = 5, the formula is $UO_2(NO_3)\cdot 5\ H_2O$, and there are 5 molecules of water of hydration in the original hydrated compound.

132. *Result:* **75.0% sulfate**

Analyze: Given a mixture of two sulfate compounds and the mass percent of one of the two compounds in the mixture, determine the mass percent of sulfate in the mixture.

Plan: Using a 100.0 gram sample of the mixture, determine the mass percent of the second compound, then calculate the mass of sulfate from each compound in the sample. Add these two masses and determine the mass percent of sulfate in the 100.0 g sample.

Execute:

100.0 g sample – 32.0 g $MgSO_4$ = 68.0 g $(NH_4)_2SO_4$

Molar mass $MgSO_4$ = 24.3050 g/mol Mg + 32.065 g/mol S + 4(15.9994 g/mol O) = 120.368 g/mol $MgSO_4$

Molar mass $(NH_4)_2SO_4$ = 2(14.0067 g/mol N) + 8(1.0079 g/mol H)

+ (32.065 g/mol S) + 4(15.9994 g/mol O) = 132.139 g/mol $(NH_4)_2SO_4$

Molar mass SO_4^{2-} = 32.065 g/mol S + 4(15.9994 g/mol O) = 96.063 g/mol SO_4^{2-}

Mass SO_4^{2-} from $MgSO_4$

$$= 32.0 \text{ g } MgSO_4 \times \frac{1 \text{ mol } MgSO_4}{120.368 \text{ g } MgSO_4} \times \frac{1 \text{ mol } SO_4^{2-}}{1 \text{ mol } MgSO_4} \times \frac{96.063 \text{ g } SO_4^{2-}}{1 \text{ mol } SO_4^{2-}} = 25.54 \text{ g } SO_4^{2-}$$

Mass SO_4^{2-} from $(NH_4)_2SO_4$

$$= 68.0 \text{ g } (NH_4)_2SO_4 \times \frac{1 \text{ mol } (NH_4)_2SO_4}{132.1392 \text{ g } (NH_4)_2SO_4} \times \frac{1 \text{ mol } SO_4^{2-}}{1 \text{ mol } (NH_4)_2SO_4} \times \frac{96.063 \text{ g } SO_4^{2-}}{1 \text{ mol } SO_4^{2-}} = 49.43 \text{ g } SO_4^{2-}$$

Total mass SO_4^{2-}= 25.54 g + 49.43 g = 74.97 g (round to one decimal place)

$$\% \ SO_4^{2-} = \frac{74.97 \text{ g } SO_4^{2-}}{100.0 \text{ g sample}} \times 100\% = 75.0\% \ SO_4^{2-} \text{ (3 sig figs)}$$

☑ *Reasonable Result Check:* The other atoms in these compounds are lightweight compared to sulfate, so it makes sense that a majority of the mass percent is sulfate.

133. *Result:* **4 Fe atoms**

Analyze and Plan: Use the percent by mass of iron to determine the mass of iron in one mole of hemoglobin(Hb), then use the molar mass of iron to determine the number of iron atoms.

Execute:
$$\frac{64458 \text{ g Hb}}{1 \text{ mol Hb}} \times \frac{0.35 \text{ g Fe}}{100. \text{ g Hb}} = 226 \text{ g Fe/mol Hb (2 sig figs)}$$

$$226 \text{ g Fe/mol Hb} \times \frac{1 \text{ mol Fe}}{55.845 \text{ g Fe}} = 4.0 \text{ mol Fe/mol Hb}$$

One mole of hemoglobin contains 4 moles of Fe. So, there must be 4 atoms of Fe in each molecule.

☑ *Reasonable Result Check:* The percent mass of iron is quit small, so it makes sense that there would only be a few iron in the formula of hemoglobin.

134. *Result:* **6.73% ^{41}K and 0.01% ^{40}K**

Analyze: Given the average atomic weight of an element, the atomic weight of three isotopes, and the percentage abundance of one isotope, determine percentage abundance of the other two isotopes.

Plan: Knowing that the weighted average of the isotope masses must be equal to the reported atomic weight, set up a relationship between the known average atomic weight and the various isotope masses using a variable to describing the other two isotope's atomic weight.

Execute: We are told that potassium has an average atomic mass of 39.0983 u and has 93.2581% ^{39}K. This percentage/100 tell us the fraction of K with this isotopes mass. Let X be the percent of the ^{40}K isotope and Y be the percent of the ^{41}K isotope.

We know that the sum of the percents must be 100%.

$$100.0000\% = \%\ ^{39}\text{K} + \%\ ^{40}\text{K} + \%\ ^{41}\text{K} = 93.2581\% + \text{X} + \text{Y}$$

We also know that the average atomic mass is the weighted average of the isotopic masses.

$$\frac{93.2581\ ^{39}\text{K atom}}{100 \text{ K atoms}} \times \left(\frac{38.963707 \text{ u}}{1 \text{ atom } ^{39}\text{K}}\right) + \frac{\text{X}\ ^{40}\text{K atom}}{100 \text{ K atoms}} \times \left(\frac{39.963999 \text{ u}}{1 \text{ atom } ^{40}\text{K}}\right) + \frac{\text{Y}\ ^{41}\text{K atom}}{100 \text{ K atoms}} \times \left(\frac{40.961825 \text{ u}}{1 \text{ atom } ^{41}\text{K}}\right)$$

$$= 39.0983 \frac{\text{u}}{\text{K atom}}$$

$$36.3368 + 0.39963999 \text{ X} + 0.40961825 \text{ Y} = 39.0983$$

Solve for the first equation for X in terms of Y:

$$\text{X} = 100.0000\% - 93.2581\% - \text{Y} = 6.7419 - \text{Y}$$

Plug this into the second equation:

$$36.3368 + 0.39963999\,(6.7419 - \text{Y}) + 0.40961825 \text{ Y} = 39.0983$$

Solve for Y:

$$2.6943 - 0.39963999\text{Y} + 0.40961825 \text{ Y} = 39.0983 - 36.3368$$

$$(0.40961825 - 0.39963999)\text{Y} = 39.0983 - 36.3368 - 2.6943$$

$$0.00997826\text{Y} = 0.0672$$

$$\text{Y} = \frac{0.0672}{0.00997826} = 6.73\%\ ^{41}\text{K}$$

Use Y to solve for X: $\text{X} = 100.0000\% - 93.2581\% - 6.73\% = 0.01\%\ ^{40}\text{K}$

So the percent abundance of the other two isotopes are **6.73% ^{41}K** and **0.01% ^{40}K**.

☑ *Reasonable Result Check:* It makes sense that these two isotopes have lower percent abundance since the average atomic weight is closer to the isotopic mass of the lightest element.

135. *Result:* **(a) Bromine monochloride (b) 114, 116, 118 (c) 116 (d) 114**

Analyze, Plan, Execute:

(a) This is a binary molecular compound, so use the naming system described in Section 2-8. BrCl is **bromine monochloride**.

(b) There are four isotopic combinations: ^{79}Br—^{35}Cl, ^{79}Br—^{37}Cl, ^{81}Br—^{35}Cl, and ^{81}Br—^{37}Cl
The sum of the mass numbers is: 79+35=114, 79+37 = 116, 81+35=116, and 81+37=118
Using the sum of the mass numbers as a criterion, there are three types: 114, 116, and 118

(c)&(d) Abundance of the diatomic molecule is determined by combining the percents
Because there are two different isotopic combinations composing the 116 type, the abundance of this type comes from two combinations ^{79}Br—^{37}Cl and ^{81}Br—^{35}Cl

^{79}Br—^{37}Cl has an abundance of (0.5069)(0.2423)=0.1228, or 12.28%
^{81}Br—^{35}Cl has an abundance of (0.4931)(0.7577)=0.3736, or 37.36%
^{79}Br—^{35}Cl has an abundance of (0.5069)(0.7577)=0.3841, or 38.41%
^{81}Br—^{37}Cl has an abundance of (0.4931)(0.2423)=0.1195, or 11.95%

The total abundance of the 116 type is 37.36% + 12.28 = 49.64% (c) The most abundant is 116 type.
The abundance of the 114 type is 38.41%. (d) The second most abundant is 114 type.
The abundance of the 118 type is 11.95%

☑ *Reasonable Result Check:* The relative abundances of the bromine isotopes are nearly equal, but there is quite a bit more ^{35}Cl than ^{37}Cl, so it makes sense to see the types that contain ^{35}Cl have greater abundance.

136. *Result:* **71.6% Fe, 18.5% Cr, and 8.96% Ni**

Analyze: Given the mass of a stainless steel sample and the masses of oxide products produced from these samples, determine the mass percent of each metal in the sample.

Plan: Use molar masses to determine moles of products. Use chemical formula to relate moles of products to moles of metal. Use molar mass of metal to calculate the mass of the metal. Divide the mass of the metal by the mass of the sample and multiply by 100% to get mass percent for each metal.

Execute: Molar mass of Fe_2O_3 = 2(55.854 g/mol Fe) + 2(15.9994 g/mol O) = 159.688 g/mol Fe_2O_3

$$10.3 \text{ g Fe}_2\text{O}_3 \times \left(\frac{1 \text{ mol Fe}_2\text{O}_3}{159.688 \text{ g Fe}_2\text{O}_3}\right) \times \left(\frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3}\right) \times \left(\frac{55.854 \text{ g Fe}}{1 \text{ mol Fe}}\right) = 7.16 \text{ g Fe}$$

$$\% \text{ Fe} = \frac{7.16 \text{ g Fe}}{10.0 \text{ g sample}} \times 100\% = 71.6\% \text{ Fe}$$

Molar mass of Cr_2O_3 = 2(51.9961 g/mol Cr) + 2(15.9994 g/mol O) = 151.9904 g/mol Cr_2O_3

$$2.71 \text{ g Cr}_2\text{O}_3 \times \left(\frac{1 \text{ mol Cr}_2\text{O}_3}{151.9904 \text{ g Cr}_2\text{O}_3}\right) \times \left(\frac{2 \text{ mol Cr}}{1 \text{ mol Cr}_2\text{O}_3}\right) \times \left(\frac{51.9961 \text{ g Cr}}{1 \text{ mol Cr}}\right) = 1.85 \text{ g Cr}$$

$$\% \text{ Cr} = \frac{1.85 \text{ g Cr}}{10.0 \text{ g sample}} \times 100\% = 18.5\% \text{ Cr}$$

Molar mass of NiO = 58.6934 g/mol Ni + 15.9994 g/mol O = 74.6928 g/mol NiO

$$1.14 \text{ g NiO} \times \left(\frac{1 \text{ mol NiO}}{74.6928 \text{ g NiO}}\right) \times \left(\frac{1 \text{ mol Ni}}{1 \text{ mol NiO}}\right) \times \left(\frac{58.6934 \text{ g Ni}}{1 \text{ mol Ni}}\right) = 0.896 \text{ g Ni}$$

$$\% \text{ Ni} = \frac{0.896 \text{ g Ni}}{10.0 \text{ g sample}} \times 100\% = 8.96\% \text{ Ni}$$

☑ *Reasonable Result Check:* The measured percentages from this sample match (in the first two sig figs) the expected percentages identified for the stainless steel used in the Gateway Arch in St. Louis, given at the beginning of the question: 71.6% ≈ 72.0%Fe, 18.5% ≈ 19.0%Cr, and 8.96% ≈ 9.0% Ni

137. *Result:* **68.50% Ga, 21.50% In, 10.0% Sn**

Analyze and Plan: There are various ways to solve this question; one is given here:
Rearrange the definition of percent by mass to use the ratios given in the question.

The mole ratio can be used to find the mass ratio using: $n_X = \dfrac{m_X}{M_X}$, where M_X = molar mass.

$$\% X = \frac{m_X}{m_X + m_Y + m_Z} \times 100\% = \frac{1}{1 + \dfrac{m_Y}{m_X} + \dfrac{m_Z}{m_X}} \times 100\%$$

Execute: We need all three mass ratios to use this equation for the three elements.

The mass ratio for Ga and In is given $\dfrac{m_{Ga}}{m_{In}} = 3.186$.

We can calculate the mass ratio for In and Sn, $\dfrac{m_{In}}{m_{Sn}}$, from the given mole ratio: $\dfrac{n_{In}}{n_{Sn}} = 2.223$

$$\frac{n_{In}}{n_{Sn}} \frac{\dfrac{m_{In}}{M_{In}}}{\dfrac{m_{Sn}}{M_{Sn}}} = \frac{m_{In} M_{Sn}}{m_{Sn} M_{In}} = 2.223$$

$$\frac{m_{In}}{m_{Sn}} = 2.223 \frac{M_{In}}{M_{Sn}} = 2.223 \frac{114.818 \text{ g/mol}}{118.710 \text{ g/mol}} = 2.150$$

We can calculate $\dfrac{m_{Ga}}{m_{Sn}}$ by combining the first two ratios: $\dfrac{m_{Ga}}{m_{Sn}} = \dfrac{m_{Ga}}{m_{In}} \dfrac{m_{In}}{m_{Sn}} = (3.186)(2.150) = 6.850$

Now, calculate the mass percent for each element:

$$\% \text{ Ga} = \frac{m_{Ga}}{m_{Ga} + m_{In} + m_{Sn}} \times 100\% = \frac{1}{1 + \dfrac{m_{In}}{m_{Ga}} + \dfrac{m_{Sn}}{m_{Ga}}} \times 100\% = \frac{1}{1 + \dfrac{1}{3.186} + \dfrac{1}{6.850}} \times 100\% = 68.50\% \text{ Ga}$$

$$\% \text{ In} = \frac{m_{In}}{m_{In} + m_{Ga} + m_{Sn}} \times 100\% = \frac{1}{1 + \dfrac{m_{Ga}}{m_{In}} + \dfrac{m_{Sn}}{m_{In}}} \times 100\% = \frac{1}{1 + 3.186 + \dfrac{1}{2.150}} \times 100\% = 21.50\% \text{ In}$$

$$\% \text{ Sn} = \frac{m_{Sn}}{m_{Sn} + m_{Ga} + m_{In}} \times 100\% = \frac{1}{1 + \dfrac{m_{Ga}}{m_{Sn}} + \dfrac{m_{In}}{m_{Sn}}} \times 100\% = \frac{1}{1 + 6.850 + 2.150} \times 100\% = 10.00\% \text{ Sn}$$

☑ *Reasonable Result Check:* The sum of the mass percents is 100%. The ratios of the percentages match the mass ratios: 68.50% Ga/21.50% In = 3.186 and 21.50% In/10.00% Sn = 2.150.

138. *Result:* **Refute the claim since 3.2 g Fe < 3.7 g Fe**

Analyze and Plan: Start with the volume of blood. Calculate the grams of hemoglobin in that amount of blood using given information. Determine the moles of hemoglobin using the molar mass. Determine the moles of Fe in the hemoglobin. Use the molar mass of Fe to find grams of Fe. Finally, compare to the mass of Fe in the iron nail.

Execute:

$$6.0 \text{ L blood} \times \left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) \times \left(\frac{15.5 \text{ g hemoglobin}}{100.0 \text{ mL blood}}\right) \times \left(\frac{1 \text{ mol hemoglobin}}{64500 \text{ g hemoglobin}}\right) \times \left(\frac{4 \text{ mol Fe}}{1 \text{ mol hemoglobin}}\right) \times \left(\frac{55.845 \text{ g Fe}}{1 \text{ mol Fe}}\right) = 3.2 \text{ g Fe}$$

3.2 g Fe < 3.7 g Fe, so while it is close to the same the mass of the nail, it is still 0.5 g less.

☑ *Reasonable Result Check:* The mass is close, so depending on how approximate the "approximately 3.7g" mass is, this claim could be considered approximately reasonable.

139. *Result:* **(a) Mg_3N_2 (b) magnesium nitride (c) fraction = 0.020**

Analyze, Plan, and Execute:

(a) The two most abundant components of the atmosphere are N_2 and O_2. While O_2 is not as abundant it is more reactive, so we will investigate the possibility that the other product is Mg_xN_y. Use the percent of Mg to determine the percent N. Calculate moles of each, set up a ratio, and find the simple whole number relationship for the formula Mg_xN_y.

$$\%N = 100.00\% - 72.25\% \text{ Mg} = 27.75\% \text{ N}$$

$$72.25 \text{ g Mg} \times \left(\frac{1 \text{ mol Mg}}{24.3050 \text{ g Mg}}\right) = 2.973 \text{ mol Mg} \qquad 27.75 \text{ g N} \times \left(\frac{1 \text{ mol N}}{14.0067 \text{ g N}}\right) = 1.981 \text{ mol N}$$

Mole ratio: 2.973 mol Mg : 1.981 mol N

Simplify by dividing by 1.981 mol: 1.500 mol Mg : 1 mol N

Simplify by multiplying by 2 3 mol Mg : 2 mol N

So, the formula is: **Mg_3N_2**

(b) This ionic compound contains ions Mg^{2+} and N^{3-}. Its name is **magnesium nitride**.

(c) Calculate how much Mg is in the MgO product and subtract from the original Mg sample mass to determine the mass of Mg in the second product. Divide this Mg mass by the sample mass to get the fraction.

$$2.512 \text{ g MgO} \times \left(\frac{1 \text{ mol MgO}}{40.03044 \text{ g MgO}}\right) \times \left(\frac{1 \text{ mol Mg}}{1 \text{ mol MgO}}\right) \times \left(\frac{24.3050 \text{ g Mg}}{1 \text{ mol Mg}}\right) = 1.515 \text{ g Mg}$$

1.546 g Mg sample – 1.515 g Mg in MgO = 0.031 g Mg in Mg_3N_2

$$\text{fraction} = \frac{0.031 \text{ g Mg}}{1.546 \text{ g sample}} = 0.020 \ (2.0\%)$$

☑ *Reasonable Result Check:* The minor product used only 2% of the Mg from the sample, which is consistent with the statement indicating that a small quantity of another magnesium compound forms.

140. *Result:* **I: K_2O II: KO_2 III: KO_3 IV: K_2O_2**

Analyze and Plan: Use mass percent K in these K_xO_y compounds to determine the percent O. Assume a 100.0 g sample, calculate the moles of K and O in the sample, set up a ratio, and find the simple whole number relationship for the empirical formula. Determine the empirical formula mass to help identify which compound has the given molar mass.

Execute:

I: 100.0% – 83.0% K = 17.0% O. A 100.0 g sample has 83.0 g K and 17.0 g O.

$$83.0 \text{ g K} \times \frac{1 \text{ mol K}}{39.0983 \text{ g K}} = 2.12 \text{ mol K} \qquad 17.0 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 1.06 \text{ mol O}$$

Mole ratio: 2.12 mol K : 1.06 mol O

Simplify by dividing by 1.06 mol: 2 mol K : 1 mol O

The empirical formula is: $\mathbf{K_2O}$

The empirical formula mass for K_2O is 2(39.0983 g/mol K) + 15.9994 g/mol O = 94.1960 g/mol.

II: 100.0% – 55.0% K = 45.0% O. A 100.0 g sample has 55.0 g K and 45.0 g O.

$$55.0\text{ g K} \times \frac{1\text{ mol K}}{39.0983\text{ g K}} = 1.41\text{ mol K} \qquad 45.0\text{ g O} \times \frac{1\text{ mol O}}{15.9994\text{ g O}} = 2.81\text{ mol O}$$

Mole ratio: 1.41 mol K : 2.81 mol O

Simplify by dividing by 1.41 mol: 1 mol K : 2 mol O

The empirical formula is: $\mathbf{KO_2}$

The empirical formula mass for KO_2 is 39.0983 g/mol K + 2(15.9994 g/mol O) = 71.0971 g/mol.

III: 100.0% – 44.9% K = 55.1% O. A 100.0 g sample has 44.9 g K and 55.1 g O.

$$44.9\text{ g K} \times \frac{1\text{ mol K}}{39.0983\text{ g K}} = 1.15\text{ mol K} \qquad 55.1\text{ g O} \times \frac{1\text{ mol O}}{15.9994\text{ g O}} = 3.44\text{ mol O}$$

Mole ratio: 1.15 mol K : 3.44 mol O

Simplify by dividing by 1.15 mol: 1 mol K : 3 mol O

The empirical formula is: $\mathbf{KO_3}$

The empirical formula mass for KO_3 is 39.0983 g/mol K + 3(15.9994 g/mol O) = 87.0965 g/mol.

IV: 100.0% –71.0% K = 29.0% O. A 100.0 g sample has 71.0 g K and 29.0 g O.

$$71.0\text{ g K} \times \frac{1\text{ mol K}}{39.0983\text{ g K}} = 1.82\text{ mol K} \qquad 29.0\text{ g O} \times \frac{1\text{ mol O}}{15.9994\text{ g O}} = 1.81\text{ mol O}$$

Mole ratio: 1.82 mol K : 1.81 mol O

Simplify by dividing by 1.81 mol: 1 mol K : 1 mol O

The empirical formula is: **KO**

The empirical formula mass for KO is 39.0983 g/mol K + 15.9994 g/mol O = 55.098 g/mol.

One of the compounds is reported to have a molar mass of 110.2 g/mol. The first three empirical formula masses are not close to an integer multiple of this but, but looking at the last compound, $(KO)_{\mathbf{n}}$:

$$n = \frac{110.2\text{ g/mol}}{55.0977\text{ g/mol}} = 2$$

Therefore, it is clear that compound IV must have a molecular formula of K_2O_2, potassium peroxide.

✓ *Reasonable Result Check:* All of the mole ratios were indisputable small integer ratios.

141. *Result:* **(a) I: XeF_4 II: XeF_2 III: XeF_6 (b) I: xenon tetrafluoride II: xenon difluoride III: xenon hexafluoride**

Analyze and Plan:

(a) Use the mass percent of Xe in compound II to determine the mass percent of F in compound II. Calculate moles of Xe and F, set up a ratio, and find the simple whole number relationship for the formula for compound II. Compound I has twice as many F atoms as Compound II (since it contains twice the mass of fluorine). Compound III has 1.5 as many F atoms as Compound I (since it contains 1.5 times the mass of fluorine contained in Compound I).

Execute:

II: 100.0% – 77.5% Xe = 22.5% F. A 100.0 g sample has 77.5 g Xe and 22.5 g F.

$$77.5 \text{ g Xe} \times \frac{1 \text{ mol Xe}}{131.293 \text{ g Xe}} = 0.590 \text{ mol Xe} \qquad 22.5 \text{ g F} \times \frac{1 \text{ mol F}}{18.9984 \text{ g F}} = 1.18 \text{ mol F}$$

Mole ratio: 0.590 mol Xe : 1.18 mol F

Simplify by dividing by 1.18 mol: 1 mol Xe : 2 mol F

The empirical formula is: $\mathbf{XeF_2}$.

I: Compound I contains twice the mass of fluorine in Compound II. That means there are twice as many atoms of F in the formula. 2 × 2 = 4. The empirical formula is: $\mathbf{XeF_4}$.

III: Compound III contains 1.5 times the mass of fluorine in Compound I. That means there are 1.5 times as many F atoms in the formula. 1.5 × 4 = 6. The empirical formula is: $\mathbf{XeF_6}$.

(b) *Explanation:* These are binary molecular compounds, so use the system described in Section 2-8.
I: The name of $\mathbf{XeF_4}$ is xenon tetrafluoride

II: The name of $\mathbf{XeF_2}$ is xenon difluoride

III: The name of $\mathbf{XeF_6}$ is xenon hexafluoride

☑ *Reasonable Result Check:* The mole ratio calculated for Compound II was very close to a small whole number. The mass of F in one mole of XeF_4 (75.9936 g) is twice the mass of F in one mole of XeF_2 (37.9968 g). The mass of F in one mole of XeF_6 (113.9904 g) is 1.5 times the mass in XeF_4 (75.9936 g).

142. *Result:* **14.7 g PCl_3 and 5.3 g PCl_5**

Analyze, Plan, and Execute:

Two compounds compose the 20.00 g sample, so $m_{\mathbf{PCl_3}} + m_{\mathbf{PCl_5}} = 20.00$ g.

$$m_{\mathbf{PCl_3}} = 20.00 \text{ g} - m_{\mathbf{PCl_5}}$$

Also, the mass chlorine in the sample can be calculated by taking 79.50% of 20.00 g:

$$m_{Cl} = 20.00 \text{ g sample} \times \left(\frac{79.50 \text{ g Cl}}{100 \text{ g sample}}\right) = 15.90 \text{ g Cl}$$

The Cl comes from two sources: $\mathbf{PCl_3}$ **and** $\mathbf{PCl_5}$.

Molar mass of PCl_3 = 30.9738 g/mol P + 3(35.453 g/mol Cl) = 137.323 g/mol PCl_3

$$\text{Mass of Cl from } PCl_3 = m_{PCl_3} \times \left(\frac{1 \text{ mol } PCl_3}{137.323 \text{ g } PCl_3}\right) \times \left(\frac{3 \text{ mol Cl}}{1 \text{ mol } PCl_3}\right) \times \left(\frac{35.453 \text{ g Cl}}{1 \text{ mol Cl}}\right) = 0.77452 m_{PCl_3}$$

Molar mass of PCl_5 = 30.9738 g/mol P + 5(35.453 g Cl) = 208.223 g/mol PCl_5

$$\text{Mass of Cl from } PCl_5 = m_{PCl_5} \times \left(\frac{1 \text{ mol } PCl_5}{208.223 \text{ g } PCl_5}\right) \times \left(\frac{5 \text{ mol Cl}}{1 \text{ mol } PCl_5}\right) \times \left(\frac{35.453 \text{ g Cl}}{1 \text{ mol Cl}}\right) = 0.851323 m_{PCl_3}$$

$$m_{\mathbf{Cl}} = = \ 0.77452\ m_{\mathbf{PCl_3}} + 0.851323 m_{\mathbf{PCl_5}}$$

Plug in $m_{\mathbf{Cl}}$ and $m_{\mathbf{PCl_3}}$ in terms of $m_{\mathbf{PCl_5}}$, from above:

$$15.90\text{g} = 0.77452(20.00 \text{ g} - m_{\mathbf{PCl_5}}) + 0.851323 m_{\mathbf{PCl_5}}$$

Solve for $m_{\mathbf{PCl_3}}$:

$$15.90 = 15.49 - 0.77452 m_{\mathbf{PCl_5}} + 0.851323 m_{\mathbf{PCl_5}}$$

$$15.90 - 15.49 = (0.851323\text{m} - 0.77452)\ m_{\mathbf{PCl_5}}$$

$$0.41 = 0.076806\ m_{\mathbf{PCl_5}}$$

$$m_{\mathbf{PCl_5}} = \frac{0.41}{0.076806} = \mathbf{5.3\ g\ PCl_5}$$

$$m_{\mathbf{PCl_3}} = 20.00\text{ g} - m_{\mathbf{PCl_5}} = 20.00\text{ g} - 5.3\text{ g } PCl_5 = \mathbf{14.7\ g\ PCl_3}$$

☑ *Reasonable Result Check:* Calculating the mass of Cl from 14.7 g PCl_3 and 5.3 g PCl_5:

$$\text{Mass of Cl from } PCl_3 = 5.3\text{ g } PCl_3 \times \left(\frac{1\text{ mol } PCl_3}{137.323\text{ g } PCl_3}\right) \times \left(\frac{3\text{ mol Cl}}{1\text{ mol } PCl_3}\right) \times \left(\frac{35.453\text{ g Cl}}{1\text{ mol Cl}}\right) = 11.4\text{ g Cl}$$

$$\text{Mass of Cl from } PCl_5 = 14.7\text{ g } PCl_5 \times \left(\frac{1\text{ mol } PCl_5}{208.223\text{ g } PCl_5}\right) \times \left(\frac{5\text{ mol Cl}}{1\text{ mol } PCl_5}\right) \times \left(\frac{35.453\text{ g Cl}}{1\text{ mol Cl}}\right) = 4.5\text{ g Cl}$$

The sum of these two masses gives $m_{\mathbf{Cl}}$ = = 11.4 g + 4.5 g = 15.9 g, which is 79.5% of 20.00 g.

143. *Result:* **This statement is verified, since 7×10^5 yr is close to $1\text{x}10^6$ yr**

Analyze, Plan, and Execute: Determine the volume of the oceans on earth.

$$\text{Surface area of the earth} = 4\pi r^2 = 4(3.1415926)\left(\frac{8000\text{ mi}}{2}\right)^2 (8000\text{mi}/2)^2 = 2 \times 10^8\text{ mi}^2$$

Surface area of the earth covered by oceans = $0.67 \times 2.0 \times 10^8\text{ mi}^2 = 1 \times 10^8\text{ mi}^2$

Volume of oceans = surface area × depth = $1\text{ mi} \times 1 \times 10^8\text{ mi}^2 = 1 \times 10^8\text{ mi}^3$

Convert mi^3 to cm^3:

$$\text{Volume} = 1\times10^8\ \text{mi}^3 \times \left(\frac{5280\text{ ft}}{1\text{ mi}}\right)^3 \times \left(\frac{12\text{ in}}{1\text{ ft}}\right)^3 \times \left(\frac{2.54\text{ cm}}{1\text{ in}}\right)^3 = 6\times10^{23}\text{ cm}^3$$

Calculate mass of seawater and mass of NaCl:

$$6\times10^{23}\ \text{cm}^3 \times \frac{1.03\ \text{g seawater}}{1\ \text{cm}^3} = 6\times10^{23}\text{ g seawater}$$

For a decrease in Mg concentration from 0.13% to 0.12%, the magnesium concentration must be reduced by 0.01%, or 0.01 g M per 100 grams of seawater.

$$6\times10^{23}\text{ g seawater} \times \left(\frac{0.01\text{ g Mg reduction}}{100\text{ g seawater}}\right) \times \left(\frac{1\text{ lb}}{453.5\text{ g Mg}}\right) \times \left(\frac{1\text{ ton}}{2000\text{ lb}}\right) \times \left(\frac{1\text{ yr}}{1.0\times10^8\text{ tons}}\right) = \mathbf{7\times10^5\ yr}$$

Considering all the approximations made here, the number 7×10^5 yr is close enough to 1×10^6 yr.

☑ *Reasonable Result Check:* The oceans have $6\times10^{23}\text{ g seawater} \times \left(\frac{0.13\text{ g Mg}}{100\text{ g seawater}}\right) = 7.8\times10^{20}$ g Mg, reducing this number to 0.12% (or 7.2×10^{20} g Mg) requires the removal of 6.0×10^{20} g Mg. A rate of 1.0×10^8 tons Mg per year removes only 3.6×10^{14} g Mg per year, so the calculated time makes sense.

144. *Result:* **5.18 g**

Analyze and Plan: Use molar mass to calculate the moles of carbon in potassium carbonate, then relate moles of carbon to moles of $K_2Zn_3[Fe(CN)_6]_2$, then use molar mass to calculate mass of $K_2Zn_3[Fe(CN)_6]_2$.

Execute:

Potassium carbonate is K_2CO_3. One mol K_2CO_3 has 1 mole C atoms. One mol $K_2Zn_3[Fe(CN)_6]_2$ has

(2 × 6 =) 12 moles of C atoms.

Molar mass K_2CO_3 = 2(39.098 g/mol K) + 12.0107 g/mol C + 3(15.9994 g/mol O) = 138.206 g/mol K_2CO_3

Molar mass $K_2Zn_3[Fe(CN)_6]_2$ = 2(39.098 g/mol K) + 3(65.38 g/mol Zn) + 2(55.845 g/mol Fe) + 12(12.0107 g/mol C) + 12(14.0067 g/mol N) = 698.235 g/mol $K_2Zn_3[Fe(CN)_6]_2$

$$12.3\text{ g }K_2CO_3 \times \left(\frac{1\text{ mol }K_2CO_3}{138.206\text{ g }K_2CO_3}\right) \times \left(\frac{1\text{ mol C}}{1\text{ mol }K_2CO_3}\right) \times \left(\frac{1\text{ mol }K_2Zn_3[Fe(CN)_6]_2}{12\text{ mol C}}\right) \times \left(\frac{698.235\text{ g }K_2Zn_3[Fe(CN)_6]_2}{1\text{ mol }K_2Zn_3[Fe(CN)_6]_2}\right) = 5.18\text{ g }K_2Zn_3[Fe(CN)_6]_2$$

☑ *Reasonable Result Check:* While the product compound has a larger molar mass, each mole of compound contains 12 C atoms, so it makes sense that the mass of product is less than the mass of the reactant.

145. *Result:* **45.0% $CaCl_2$**

Analyze and Plan: A mixture contains calcium chloride and sodium chloride. Use molar mass to calculate the moles of calcium in CaO, then relate moles of calcium to moles of $CaCl_2$, then use molar mass to calculate mass of $CaCl_2$.

Execute: Calcium chloride is $CaCl_2$ and sodium chloride is NaCl. The calcium in $CaCl_2$ is converted to Calcium carbonate, $CaCO_3$, which is then converted to calcium oxide, CaO. One mol $CaCl_2$ has 1 mol Ca atoms. One mol CaO has 1 mol Ca atoms.

Molar mass of $CaCl_2$ = 40.078 g/mol Ca + 2(35.453 g/mol Cl) = 110.984 g/mol $CaCl_2$

Molar mass of CaO = 40.078 g/mol Ca + 15.9994 g/mol O = 56.077 g/mol CaO

$$0.959\text{ g CaO} \times \left(\frac{1\text{ mol CaO}}{56.077\text{ g CaO}}\right) \times \left(\frac{1\text{ mol Ca}}{1\text{ mol CaO}}\right) \times \left(\frac{1\text{ mol }CaCl_2}{1\text{ mol Ca}}\right) \times \left(\frac{110.984\text{ g }CaCl_2}{1\text{ mol }CaCl_2}\right) = 1.90\text{ g }CaCl_2$$

$$\%\ CaCl_2\text{ in mix } = \frac{1.90\text{ g }CaCl_2}{4.22\text{ g mixture}} \times 100\% = 45.0\ \%\ \ CaCl_2$$

☑ *Reasonable Result Check:* The mass of the $CaCl_2$ is smaller than the sample mass.

146. *Result:* **58.0% M in MO**

Analyze: Given the mass percent of one compound, M_2O, containing one known element, O, and one unknown element, M, calculate the percent by mass of another compound, MO.

Plan: Choose a convenient sample mass of M_2O, such as 100.0 g. Find the mass of M and O in the sample, using the given mass percent. Using the molar mass of oxygen as a conversion factor, determine the number of moles of oxygen, then use the formula stoichiometry of M_2O as a conversion factor to determine the number of moles of M. Find the molar mass of M by dividing the mass of M by the moles of M. Use the molar mass of M, and the formula stoichiometry of MO, to determine the mass percent of M in MO.

Execute: 73.4% M in M_2O means that 100.0 grams of M_2O contains 73.4 grams of M.

Mass of O = 100.0 g M_2O – 73.4 g M = 26.6 g O

Formula stoichiometry: 1 mol of M_2O contains 2 mol M and 1 mol O.

$$26.6\text{ g O} \times \frac{1\text{ mol O}}{15.9994\text{ g O}} \times \frac{2\text{ mol M}}{1\text{ mol O}} = 3.33\text{ mol M}$$

$$\text{Molar mass of M} = \frac{\text{mass of M in sample}}{\text{mol of M in sample}} = \frac{73.4\text{ g M}}{3.33\text{ mol M}} = 22.1\frac{\text{g}}{\text{mol}}$$

Molar mass of MO = 22.1 g/mol M + 15.9994 g/mol O = 38.07 g/mol MO

$$\%\,M = \frac{\text{mass of M / mol MO}}{\text{mass of MO/ mol MO}} \times 100\% = \frac{22.1\ \text{g M}}{38.07\ \text{g MO}} \times 100\% = 58.0\ \%\ \text{M in MO}$$

☑ *Reasonable Result Check:* It makes sense that the compound with more atoms of M has a higher mass percent of M. The closest element to M's atomic mass (22.1) is sodium (atomic mass = 22.99). If M is sodium, the two compounds would probably be sodium oxide (Na_2O) and sodium peroxide (Na_2O_2, a compound made up of two Na^+ ions and one O_2^{2-} ion. The simple ratio of Na and O atoms in this compound is 1:1).

147. *Result:* **(a) –2 (b) Al_2X_3 (c) Se**

Analyze, Plan, and Execute:

(a) A group 6A element is likely to have a –2 charge, since: anion charge = (group number) – 8 = 6 – 8 = –2

(b) Aluminum ion, Al^{3+} combines with X^{2-} to form Al_2X_3.

(c) Given the percent by mass of an element in a compound and the compound's formula including an unknown element, determine the identity of the unknown element.

Choose a convenient sample of Al_2X_3, such as 100.00 g. Using the percent by mass, determine the number of grams of Al and X in the sample. Use the molar mass of Al as a conversion factor to get the moles of Al. Use the formula stoichiometry as a conversion factor to get the moles of X. Determine the molar mass by dividing the grams of X in the sample, by the moles of X in the sample. Using the periodic table, determine which Group 6A element has a molar mass nearest this value.

The compound is 18.55% Al by mass. This means that 100.00 g of Al_2X_3 contains 18.55 grams Al and the rest of the mass is from X.

Mass of X in sample = 100.00 g Al_2X_3 – 18.55 g Al = 81.45 g X

Formula stoichiometry: 1 mol of Al_2X_3 contains 2 mol of Al atoms.

$$18.55\ \text{g Al} \times \frac{1\ \text{mol Al}}{26.9815\ \text{g Al}} \times \frac{3\ \text{mol X}}{2\ \text{mol Al}} = 1.031\ \text{mol X}$$

$$\text{Molar Mass of X} = \frac{\text{mass of X in sample}}{\text{moles of X in sample}} = \frac{81.45\ \text{g X}}{1.031\ \text{mol X}} = 78.98\ \text{g/mol}$$

The periodic table indicates that X = Se (Z = 34, with atomic weight = 78.96 g/mol).

☑ *Reasonable Result Check:* The molar mass calculated is close to that of the Group 6A element, Se. Ions of the elements in Group 6A will typically have 2– charge, and the formula would be Al_2Se_3.

Conceptual Challenge Problems

Many of the Conceptual Challenge Problems in this book go beyond the topic in the chapter, and can help to stretch and integrate students' understanding of how the subjects in various chapters interconnect. Many of them are open-ended and philosophical and as such can be used for group work, special projects, or class discussions. It is not always possible to provide a complete "solution" for some of these questions; however, some additional information and helpful instruction is provided for the use and evaluation of each question.

CP2.A The ratio of the stacks should give whole number proportions.

$$\text{ratio}_{2:1} = \frac{15.96\ \text{g}}{9.12\ \text{g}} = 1.75 = \frac{7}{4} \qquad \text{ratio}_{3:1} = \frac{27.36\ \text{g}}{9.12\ \text{g}} = 3.00 = \frac{3}{1} \qquad \text{ratio}_{3:2} = \frac{27.36\ \text{g}}{15.96\ \text{g}} = 1.714 = \frac{12}{7}$$

$$\text{If Stack 1 has four dimes: mass of one dime} = \frac{9.12\ \text{g}}{4} = 2.280\ \text{g}$$

$$\text{If Stack 2 has seven dimes: mass of one dime} = \frac{15.96\ \text{g}}{7} = 2.280\ \text{g}$$

If Stack 3 has 12 dimes: $\text{mass of one dime} = \dfrac{27.36\text{ g}}{12} = 2.280\text{ g}$

The proposed mass of the dime is 2.280 g (or a mass that is 2.28 divided by some integer, (such as 1.14 = 2.28/2). All three masses are accounted for in this description.

CP2.B $18\text{ billion years} \times \dfrac{1{,}000{,}000{,}000\text{ years}}{1\text{ billion years}} \times \dfrac{365.25\text{ days}}{1\text{ year}} \times \dfrac{24\text{ h}}{1\text{ day}} \times \dfrac{3600\text{ s}}{1\text{ h}} = 5.7 \times 10^{17}\text{ s}$

$$5.7 \times 10^{17}\text{ C atoms} \times \frac{1\text{ mole C}}{6.022 \times 10^{23}\text{ C atoms}} \times \frac{12.0107\text{ g C}}{1\text{ mole C}} = 0.000011\text{ g C}$$

This mass it too small to be detected on a balance that measures masses as small as ±0.0001 g.

CP2.C The students should be asked to look at the numbers and contemplate what they can determine about them, before picking up a calculator. The three compounds have different formulas, as manifested most clearly by the % Ey differences. The second of these has a larger number of Ey atoms in the formula, due to its larger% mass; the third has the least Ey. The best way to quantitatively compare these compounds is to scale the samples so that they have the same masses of one element. For example:

100.00 g of compound B has 40.002 g Ex, 6.7142 g Ey, and 53.284 g Ez. Scaling sample A by a factor of 1.0671 (calculated by dividing the $\%Ex_B/\%Ey_A$) shows that 106.71 grams of A has 40.002 g Ex, 13.427 g Ey, and 53.284 g Ez. Scaling sample C by a factor of 0.983212 (calculated by dividing the $\%Ex_B/\%Ey_C$) shows that 98.3212 grams of C has 40.002 g Ex, 5.0356 g Ey, and 53.283 g Ez.

A glance at these scaled masses shows that these three compounds have the same proportion of Ex to Ez. So the ratio of Ex atoms to Ez atoms in each of these formulas is the same. The ratio of Ey in each of them is: A/B = 2/1 and B/C = 1.333 = 4/3. Comparing this information, it is possible to determine that the Ey ratio in the three compounds is A:B:C = 8:4:3.

CP2.D Using what is known about the relationship between Ex and Ez in the given formula, it is possible to determine the coefficient for these two atoms in the other two formulas (since they must be the same). The 8:4:3 ratio found in the first problem allows for the relative determination of the Ey element in the other two formulas.

Compound A	**Compound B**	**Compound C**
$\mathbf{ExEy_4Ez}$	$ExEy_2Ez$	$\mathbf{ExEy_{3/2}Ez}$ **(?)**
$Ex_6Ey_8Ez_3$	$\mathbf{Ex_6Ey_4Ez_3}$	$\mathbf{Ex_6Ey_3Ez_3}$
$\mathbf{Ex_3Ey_{16/3}Ez}$	$\mathbf{Ex_3Ey_{8/3}Ez}$	Ex_3Ey_2Ez
$\mathbf{Ex_9Ey_4Ez_6}$	$Ex_9Ey_2Ez_6$	$\mathbf{Ex_9Ey_{3/2}Ez_6}$ **(?)**
$\mathbf{ExEy_{16/3}Ez_3}$ **(?)**	$\mathbf{ExEy_{8/3}Ez_3}$ **(?)**	$ExEy_2Ez_3$
$Ex_3Ey_8Ez_3$	$\mathbf{Ex_3Ey_4Ez_3}$	$\mathbf{Ex_3Ey_3Ez_3}$

CP2.E (a) If the mass of Ez is 1.3320 times heavier than the mass of Ex, then the atoms must be present in equal proportion, since the mass ratio in the scaled samples is the same (54.284 g/40.002 g), indicating that the formula must be $Ex_nEy_{8n}Ez_n$.

(b) If the mass of Ex is 11.916 times heavier than the mass of Ey, then there must be 1:2 atom ratio of Ex to Ey in compound B, where the mass ratio is 11.916/2.

(c) If the mass ratios are known, then the formulas can be determined.