4.1 PROPORTIONS IN COMPOUNDS

PRACTICE

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Understanding Concepts

1. Students will largely reproduce Table 1 on page 160 of the text, adding information on the chemical properties of the two compounds, e.g.:

Comparing CO and CO_2

	CO	CO_2
physical properties		
melting point (°C)	–199.0	–78.5
boiling point (°C)	–191.5	–78.5
density (g/L)	1.250	1.977
colour	none	none
solubility (g/100mL H_2O)	0.004	0.339
chemical properties		
odour	none	none
toxicity	high	low
taste	none	slightly acid
blood	binds to hemoglobin	affects acidity
photosynthesis	not used	reactant
applications	?	dry ice; carbonation; fire extinguishers

2. Carbon monoxide is toxic because it bonds to hemoglobin in the bloodstream, preventing blood cells from carrying oxygen to body tissues.
3. Carbonated beverages are made with carbon dioxide gas. Its physical property of dissolving under pressure and escaping when pressure is low makes the drink "fizzy." Its chemical properties make the drink taste "tangy."
   Carbon dioxide is useful in fire extinguishers. It can be liquified and stored under pressure, and will propel itself out of its container when opened. It is chemically unreactive and doesn't support combustion of most substances.

PRACTICE

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Understanding Concepts

4. When compounds form from elements, the ratio of masses of those elements in the compound is always precisely the same.
5. Stoichiometry refers to the study of quantities of substances involved in chemical reactions.
6. (a) Nonmetals can often form many bonds, and bond to each other in several different ways, resulting in numerous different compounds.
   (b) Sulfur and phosphorus should bond to oxygen in more than one way, since they are nonmetals.
7. (a) $CuCl_{(s)}$ and $CuCl_{2(s)}$
   (b) $FeO_{(s)}$ and $Fe_2O_{3(s)}$
   (c) $PbS_{(s)}$ and $PbS_{2(s)}$

Making Connections

8. Examples might include medical lab technologist, food chemistry, cosmetics chemistry, plastics chemistry, or detergent/cleanser chemistry analyst, and so on.

SECTION 4.1 QUESTIONS

(Page 162)

Understanding Concepts

1. Statements (a), (b), and (c) illustrate the law of definite proportions, and (d) does not.

Applying Inquiry Skills

2. **Experimental Design**

   Aqueous silver ions and chromate ions will be mixed indifferent ratios to see how they combine. The solutions of ions will each contain the same number of ions per unit volume.

Materials

- lab apron
- eye protection
- 5 small test tubes of equal size
- test-tube rack
- eyedropper
- 10 mL of silver nitrate solution (0.10 mol/L)
- 10 mL of sodium chromate solution (0.10 mol/L)
- distilled water

Procedure

Safety: Chromate compounds are poisonous if ingested. Wash hands after use.

1. Number the test tubes from 1 to 5, and place in the test-tube rack.
2. Using the dropper, add drops of silver nitrate solution to each test tube. Add 2, 4, 6, 8, and 10 drops respectively, to tubes 1, 2, 3, 4, and 5.
3. Wash the dropper thoroughly with distilled water and use the same dropper to add drops of sodium chromate solution to each test tube. Add 10, 8, 6, 4, and 2 drops respectively, to tubes 1, 2, 3, 4, and 5. After you have finished putting drops in each tube, the test tubes should be filled to equal depth since they contain the same number of drops (12 drops total).
4. Swirl each test tube gently to mix the contents. Allow the precipitates to settle for about 5 minutes.
5. Wash your hands thoroughly.

Analysis

(a) The test tube with the most precipitate indicates the ratio of combination of silver and chromate ions — which will be proportional to the ratio of drops used in that tube.

(b) This experiment can be done with any desired ratio of drops, but the result will always be that silver and chromate combine in a 2:1 ratio, showing that the proportions of the precipitate compound are fixed.

Making Connections

3.

Compound	Colour	Hazard	Use
$N_2O_{(g)}$	colourless	slight	anesthetic
$NO_{(g)}$	colourless	toxic/irritant	bleaching
$NO_{2(g)}$	red–brown	very toxic	rocket fuels

GO TO www.science.nelson.com, Chemistry 11, Teacher Centre.

4.2 RELATIVE ATOMIC MASS AND ISOTOPIC ABUNDANCE

PRACTICE

(Page 164)

Understanding Concepts

1. Atomic mass, like any other quantity, has to have a reference value with which it can be compared, for accuracy.
2. The atomic mass unit is exactly 1/12 the mass of a single atom of carbon-12. The unit symbol is u.
3. Hydrogen, oxygen, and carbon atoms have been used as reference elements for atomic mass.
4. $1.92 \times 12\ \text{u} = 23.0\ \text{u}$ is the relative atomic mass of sodium.

Making Connections

5. A communication system must be convenient, simple, and practical. To be international it must also be accepted universally. The SI and IUPAC systems are both designed around these criteria.

PRACTICE

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Understanding Concepts

6. (a) Average atomic mass is the average value of the mass of atoms in a sample of a naturally occurring element.
   (b) Isotopes of an element are atoms with the same number of protons, but different numbers of neutrons, and so different masses.
   (c) Isotopic abundance refers to the proportion of atoms of an element that are specific isotopes, usually given as a percentage.
7. The isotope C-12 refers to atoms with 6 protons and 6 neutrons.
8. The atomic mass of carbon is not exactly 12 u because a small percentage of carbon atoms are C-13 atoms, which increase the average value of the mass.
9. Since atoms exist as indivisible objects for chemistry purposes, we avoid talking about fractions of atoms to avoid confusion.
10. Assume 10 000 K atoms, for convenience
    93.10% or 9310 atoms are K-39
    6.90% or 690 atoms are K-41

$$m_{tot} = (9310 \times 39\ \text{u}) + (690 \times 41\ \text{u})$$

$$m_{tot} = 391\ 380\ \text{u}$$

$$m_{av} = \frac{391\ 380\ \text{u}}{10\ 000}$$

$$m_{av} = 39.14\ \text{u}$$

    The actual average atomic mass value for potassium is 39.10 u, which is very close to the value calculated by this method.
11. Assume 10 000 Ar atoms, for convenience
    99.60% or 9960 atoms are Ar-40
    0.34% or 34 atoms are Ar-36
    0.06% or 6 atoms are Ar-38

$$m_{tot} = (9960 \times 40\ \text{u}) + (34 \times 36\ \text{u}) + (6 \times 38\ \text{u})$$

$$m_{tot} = 399\ 852\ \text{u}$$

$$m_{av} = \frac{399\ 852\ \text{u}}{10\ 000}$$

$$m_{av} = 39.99\ \text{u}$$

    The actual average atomic mass value for argon is 39.95 u, which is very close to the value calculated by this method.

SECTION 4.2 QUESTIONS

(Page 167)

Understanding Concepts

1. 9×12 u = 108 u (assume *both* numbers are exact), which is very close to the average value for silver, Ag, which is actually 107.87 u.
2. The atomic mass unit is defined as 1/12 the mass of a C-12 atom. Atomic mass is the actual average mass of the atoms of a given element.
3. Because the elements are compared to each other based on the way they combine when they react, the reaction proportions must be known.
4. Chlorine atoms have an average atomic mass of 35.45 u, because chlorine is made up of about 75% Cl-35 atoms, and about 25% Cl-37 atoms.
5. Assume 1000 B atoms, for convenience
   19.8% or 198 atoms are B-10
   80.2% or 802 atoms are B-11

$$m_{tot} = (198 \times 10\ \text{u}) + (802 \times 11\ \text{u})$$

$$m_{tot} = 10\,802\ \text{u}$$

$$m_{av} = \frac{10\,802\ \text{u}}{1000}$$

$$m_{av} = 10.8\ \text{u}$$

   The average atomic mass for boron is 10.8 u, calculated by this method.

4.3 THE MOLE AND MOLAR MASS

PRACTICE

(Page 168)

Understanding Concepts

1. A mole of anything is the same number of things as the number of C atoms there would be in exactly 12 g of the isotope C-12.
2. Avogadro's constant is 6.02×10^{23} (rounded to three digits ...).
   Note: Since the concepts of "pure" C-12 and "exactly" 12 g are imaginary, there is no pretense in the scientific community that we will ever know the "exact" value for Avogadro's constant. The mole is a purely theoretical definition. As technology improves, we are, of course, able to determine the value to greater precision. Rounded to six digits, the precision routinely stated in postsecondary level work, the currently accepted value is $6.022\,14 \times 10^{23}$. The Canadian Metric Practice Guide lists 8 digits — $6.022\,136\,7 \times 10^{23}$. The most precise recent reported value, obtained from ion X-ray diffraction evidence, is $6.022\,141\,99 \times 10^{23}$. This constant, like many others that are frequently used, is usually rounded to three digits for high-school calculations.
3. 602 000 000 000 000 000 000 000 (only the first 3 digits are significant)
   Note: It is useful for students to extend this number to see how large it is, being mindful that all those trailing zeros represent numbers they don't know. It is also thought-provoking for students to consider that when written in scientific notation to the usual high-school text precision, the error of the number is 2×10^{20}, meaning we blithely ignore some two hundred millions of millions of millions, as being too small to make a noticeable difference.
4. $$N_{CO_2} = 3.00\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{CO_2} = 1.81 \times 10^{24}\ \text{molecules}$$

   There are 1.81×10^{24} molecules of carbon dioxide in the sample.

5. $N_{Ar} = 0.500 \cancel{mol} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \cancel{mol}}$

$N_{Ar} = 3.01 \times 10^{23}$ atoms

There are 3.01×10^{23} atoms of argon in the sample.

6. (a) $m = \frac{1.43 \text{ kg}}{12 \text{ oranges}}$ (12 is an exact value — counted)

$m = 0.119$ kg/orange = 119 g/orange

The average mass of one of these oranges is 119 g.

(b) $m = \frac{1.01 \text{ g}}{1 \cancel{mol}} \times \frac{1 \cancel{mol}}{6.02 \times 10^{23} \text{ H atoms}}$

$m = 1.68 \times 10^{-24}$ g/H atom (average value)

The average mass of an atom in a hydrogen sample is 1.68×10^{-24} g.

Note: No actual hydrogen atom, of course, really has this mass value, since the relative atomic molar mass used to derive it is an *average* value, itself derived from the natural isotopic composition of the element.

PRACTICE

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Understanding Concepts

7. Molar mass is the mass of one mole (the Avogadro number) of any entity, usually atoms, molecules, ions, or ionic compound formula units. The SI unit is technically kg/mol, but is commonly used and stated as g/mol.

8. 1 mol $Ca(OH)_2$ = 1 mol Ca + 2 mol O + 2 mol H

$M_{Ca(OH)_2} = [(40.08 \times 1) + (16.00 \times 2) + (1.01 \times 2)]$ g/mol

$M_{Ca(OH)_2} = 74.10$ g/mol

The molar mass of calcium hydroxide is 74.10 g/mol.

9. 1 mol Cl_2 = 2 mol Cl

$M_{Cl_2} = (35.45 \times 2)$ g/mol

$M_{Cl_2} = 70.90$ g/mol

The molar mass of chlorine is 70.90 g/mol

10. 1 mol OH^- = 1 mol O + 1 mol H + 1 mol e^-

$M_{OH^-} = [(16.00 \times 1) + (1.01 \times 1) + (\text{negligible mass})]$ g/mol

$M_{OH^-} = 17.01$ g/mol

The molar mass of hydroxide ions is 17.01 g/mol.

Note: Students are taught routinely to simply ignore the electrical charges of ions when calculating molar masses, because the tiny variation in mass due to electron increase/decrease is negligible, never enough to affect the value.

11. The element is most likely gold, $Au_{(s)}$, which has a molar mass listed on the periodic table of elements of 196.97 g/mol.

12. $m = 8.0 \cancel{mol} \times \frac{67.2 \text{ g}}{\cancel{mol}}$

$m = 5.4 \times 10^2$ g = 0.54 kg

The mass of the substance is 0.54 kg.

13. (a) Diatomic means composed of two-atom molecules.
(b) Hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine form diatomic molecules.
(c) Avogadro's number of molecules (6.02×10^{23} molecules) are present.

SECTION 4.3 QUESTIONS

(Page 171)

Understanding Concepts

1. Avogadro's constant is the number of atoms of carbon in exactly 12 g of the isotope C-12, also called one mole. It is useful in chemistry because a mole is defined so that the mass of this amount in grams will equal numerically the relative atomic mass of any atom.
2. Atomic mass is the average value of the mass of atoms of a particular element. Molar mass is the mass of one mole of any entity.
3. 1 mol $C_{12}H_{22}O_{11}$ = 12 mol C + 22 mol H + 11 mol O

   $M_{C_{12}H_{22}O_{11}} = [(12.01 \times 12) + (1.01 \times 22) + (16.00 \times 11)]$ g/mol

   $M_{C_{12}H_{22}O_{11}} = 342.34$ g/mol

   $N_A = 6.02 \times 10^{23}$ entities/mol

   $m_{C_{12}H_{22}O_{11}} = ?$

   $$m = \frac{342.34 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{6.02 \times 10^{23} \text{ molecules}}$$

   $m = 5.69 \times 10^{-22}$ g/molecule (on average)

   The average mass of a molecule in a sucrose sample is 5.69×10^{-22} g.

   Note: No actual single sucrose molecule, of course, has this mass value. It is an average value derived from the natural isotope mixtures of elements in this compound.

4. 1 mol C_8H_{18} = 8 mol C + 18 mol H

   $M_{C_8H_{18}} = [(12.01 \times 8) + (1.01 \times 18)]$ g/mol

   $M_{C_8H_{18}} = 114.26$ g/mol

   The molar mass of octane is 114.26 g/mol.

   Note: You may find it convenient, from this point on in the course, to assume that students will calculate molar masses accurately, i.e., not requiring them to show the addition of periodic table values. Since all of the periodic table (average atomic molar mass) values are given to two decimal places, and the precision rule for addition always applies, any molar mass calculation will automatically be precise to hundredths of a gram per mole.

Making Connections

5. Molar masses:

$H_{2(g)}$	2.02 g/mol
$He_{(g)}$	4.00 g/mol
$N_{2(g)}$	28.02 g/mol
$O_{2(g)}$	16.00 g/mol
$CO_{2(g)}$	44.01 g/mol

   The density of each gas is found as follows; assume the volume of a mole of each is approximately 22.4 L.

   $$\text{density}_{H_2} = \frac{2.02 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

   $\text{density}_{H_2} = 0.0902$ g/L

   The density of hydrogen gas at STP is 0.0902 g/L.

   $$\text{density}_{He} = \frac{4.00 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{He} = 0.179 \text{ g/L}$

The density of helium gas at STP is 0.179 g/L.

$$\text{density}_{N_2} = \frac{28.02 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{N_2} = 1.25 \text{ g/L}$

The density of nitrogen gas at STP is 1.25 g/L.

$$\text{density}_{O_2} = \frac{32.00 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{O_2} = 1.43 \text{ g/L}$

The density of oxygen gas at STP is 1.43 g/L.

$$\text{density}_{CO_2} = \frac{44.01 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{CO_2} = 1.96 \text{ g/L}$

The density of carbon dioxide gas at STP is 1.96 g/L.

Note : The molar volume of 22.4 L/mol at STP for gases is an approximation derived from a theoretical "ideal" gas system, rounded to three significant digits. See page 469 of the text (Did You Know?) for more precise values for some real gases.

Reflecting

6. Molar mass is a conversion factor that can be used to convert measured masses of substances into numerical amounts, in moles. This will be useful because the numbers in chemical reaction equations represent numerical values.

4.4 CALCULATIONS INVOLVING THE MOLE CONCEPT

PRACTICE

(Page 172)

Understanding Concepts

1. $m_{NaCl} = 2.5 \text{ g}$

$M_{NaCl} = 58.44 \text{ g/mol}$

$n_{NaCl} = ?$

$$n_{NaCl} = 2.5 \cancel{\text{g}} \times \frac{1 \text{ mol}}{58.44 \cancel{\text{g}}}$$

$$n_{NaCl} = 0.043 \text{ mol} = 43 \text{ mmol}$$

The amount of sodium chloride is 43 mmol.

2. $m_{C_6H_{12}O_6} = 1.0 \text{ kg}$

$M_{C_6H_{12}O_6} = 180.18 \text{ g/mol}$

$n_{C_6H_{12}O_6} = ?$

$$n_{C_6H_{12}O_6} = 1.0 \text{ k}\cancel{\text{g}} \times \frac{1 \text{ mol}}{180.18 \cancel{\text{g}}}$$

$$n_{C_6H_{12}O_6} = 0.056 \text{ kmol} = 5.6 \text{ mol}$$

The amount of glucose is 5.6 mol.

3. m_{O_2} = 25.0 g

M_{O_2} = 32.00 g/mol

n_{O_2} = ?

$$n_{O_2} = 25.0\ \cancel{g} \times \frac{1\ \text{mol}}{32.00\ \cancel{g}}$$

$$n_{O_2} = 0.781\ \text{mol}$$

The amount of oxygen is 0.781 mol (or 781 mmol).

4. (a) m_C = 24.0 g

M_C = 12.01 g/mol

n_C = ?

$$n_C = 24.0\ \cancel{g} \times \frac{1\ \text{mol}}{12.01\ \cancel{g}}$$

$$n_C = 2.00\ \text{mol}$$

The amount of carbon atoms in the compound is 2.00 mol.

m_H = 6.0 g

M_H = 1.01 g/mol

n_H = ?

$$n_H = 6.0\ \cancel{g} \times \frac{1\ \text{mol}}{1.01\ \cancel{g}}$$

$$n_H = 5.9\ \text{mol}$$

The amount of hydrogen atoms in the compound is 5.9 mol.

m_O = 16.0 g

M_O = 16.00 g/mol

n_O = ?

$$n_O = 16.0\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$

$$n_O = 1.00\ \text{mol}$$

The amount of oxygen atoms in the compound is 1.00 mol.

(b) The mole ratio of C:H:O is approximately 2:6:1.

(c) The likely formula for the compound then, is $C_2H_6O_{(l)}$.

PRACTICE

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Understanding Concepts

5. $n_{Mg(OH)_2}$ = 0.45 mol

$M_{Mg(OH)_2}$ = 58.33 g/mol

$m_{Mg(OH)_2}$ = ?

$$m_{Mg(OH)_2} = 0.45\ \cancel{\text{mol}} \times \frac{58.33\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$m_{Mg(OH)_2} = 26\ \text{g}$$

The mass of magnesium hydroxide is 26 g.

6. n_{NH_3} = 87 mmol

M_{NH_3} = 17.04 g/mol

m_{NH_3} = ?

$$m_{NH_3} = 87\ \text{m}\cancel{\text{mol}} \times \frac{17.04\ \text{g}}{1\ \cancel{\text{mol}}}$$
$$m_{NH_3} = 1.5 \times 10^3\ \text{mg} = 1.5\ \text{g}$$

The mass of ammonia is 1.5 g.

7. $n_{C_8H_6O_4} = 63.28$ mol

$M_{C_8H_6O_4} = 166.14$ g/mol

$m_{C_8H_6O_4} = ?$

$$m_{C_8H_6O_4} = 63.28\ \cancel{\text{mol}} \times \frac{166.14\ \text{g}}{1\ \cancel{\text{mol}}}$$
$$m_{C_8H_6O_4} = 10.51 \times 10^3\ \text{g} = 10.51\ \text{kg}$$

The mass of 1,4-benzenediotic acid is 10.51 kg.

8. $n_{HC_9H_7O_4} = 1.0$ mmol

$M_{HC_9H_7O_4} = 180.17$ g/mol

$m_{HC_9H_7O_4} = ?$

$$m_{HC_9H_7O_4} = 1.0\ \text{m}\cancel{\text{mol}} \times \frac{180.17\ \text{g}}{1\ \cancel{\text{mol}}}$$
$$m_{HC_9H_7O_4} = 1.8 \times 10^2\ \text{mg} = 0.18\ \text{g}$$

The mass of acetylsalicylic acid (Aspirin) is 0.18 g.

Try This Activity : Counting Atoms, Molecules, and Other Entities

(Page 176)

Calculate the evaporation rate (in molecules) of a drop of water:
mass of 50 drops of water = 1.95 g
mass of 1 drop of water = 1.95 g/50 = 0.039 g
time of 1 drop of water to evaporate = 65 min = 3.9×10^3 s

$$n_{H_2O} = 0.039\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{18.02\ \cancel{\text{g}}}$$
$$= 0.00216\ \text{mol}$$

There is 0.00216 mol of H_2O molecules in every drop.

$$N_{H_2O} = 0.00216\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$
$$= 1.30 \times 10^{21}\ \text{molecules}$$

There are 1.30×10^{21} molecules of H_2O in every drop.

$$\text{evaporation rate} = \frac{1.30 \times 10^{21}\ \text{molecules}}{3.9 \times 10^3\ \text{s}}$$

$$\text{evaporation rate} = 3.34 \times 10^{17}\ \text{molecules/s}$$

Water molecules evaporate at the average rate of 3.34×10^{17} molecules/s.

Calculate the number of atoms and the value of each atom in a copper penny:
mass of 1 penny = 2.44 g

$$n_{Cu} = 2.44\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{63.55\ \cancel{\text{g}}}$$
$$= .0384\ \text{mol}$$
$$N_{Cu} = .0384\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \cancel{\text{mol}}}$$
$$N_{Cu} = 2.31 \times 10^{22}\ \text{atoms}$$

There are 2.31×10^{22} atoms in a copper penny.

$$\text{value of a copper atom} = \frac{1 \text{ cent}}{2.31 \times 10^{22} \text{ atoms}}$$

$$\text{value of a copper atom} = 4.33 \times 10^{-23} \text{ cents}$$

Each copper atom in a penny has a value of 4.33×10^{-23} cents.

Measure into a graduated cylinder half a mole of sucrose molecules:

$M_{C12H22O11} = 337.34$ g/mol

$$\text{mass of half a mole of sucrose} = \frac{337.34 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{2}$$

$$= 168.67 \text{ g}$$

When poured into a graduated cylinder, the volume of sucrose = 185 mL.

Measure into a graduated cylinder the quantity of sugar that contains two moles of carbon atoms:

One mole of sucrose, $C_{12}H_{22}O_{11(s)}$, contains twelve moles of carbon atoms. Therefore, you would need one-sixth of a mole of sucrose to get two moles of carbon atoms.

$$\text{required mass of sucrose} = \frac{337.34 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{6}$$

$$= 56.22 \text{ g}$$

When poured into a graduated cylinder, the volume of sucrose = 60 mL.

Calculate the number of atoms needed to write your name in chalk:

initial mass of chalk = 14.23 g

final mass of chalk = 14.15 g

change in mass of chalk = 0.08 g

$M_{CaCO3} = 100.09$ g/mol

number of atoms in each formula unit = 5

$$n_{CaCO_3} = 0.08 \cancel{\text{g}} \times \frac{1 \text{ mol}}{100.09 \cancel{\text{g}}}$$

$$= 0.000799 \text{ mol}$$

$$N_{CaCO_3} = 0.000799 \cancel{\text{mol}} \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \cancel{\text{mol}}}$$

$$= 4.81 \times 10^{20} \text{ formula units}$$

$$\text{number of atoms} = 4.81 \times 10^{20} \cancel{\text{form unit}} \times \frac{5 \text{ atoms}}{1 \cancel{\text{form unit}}}$$

$$= 2.40 \times 10^{21} \text{ atoms}$$

The number of atoms needed to write my name in chalk is 2.40×10^{21}.

Calculate the number of sodium ions in a solution of 3.00 g of NaCl in 200 mL of water:

$m_{NaCl} = 3.00$ g

$M_{NaCl} = 58.44$ g/mol

$$N_{NaCl} = 3.00 \cancel{\text{g}} \times \frac{1 \cancel{\text{mol}}}{58.44 \cancel{\text{g}}} \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \cancel{\text{mol}}}$$

$$= 3.09 \times 10^{22} \text{ formula units}$$

Since there is one sodium ion in every NaCl formula unit, the number of sodium ions is 3.09×10^{22}.

Calculate the number of iron atoms in a nail:

$m_{nail} = 4.16$ g

$M_{Fe} = 55.85$ g/mol

$$N_{Fe} = 4.16 \cancel{\text{g}} \times \frac{1 \cancel{\text{mol}}}{55.85 \cancel{\text{g}}} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \cancel{\text{mol}}}$$

$$= 4.48 \times 10^{22}$$

There are 4.48×10^{22} atoms in the iron nail.

Calculate the number of years to span a mole of seconds:

$$\text{Number of seconds in one year} = \frac{60 \text{ s}}{1 \cancel{\text{min}}} \times \frac{60 \cancel{\text{min}}}{1 \cancel{\text{h}}} \times \frac{24 \cancel{\text{h}}}{1 \cancel{\text{d}}} \times \frac{365 \cancel{\text{d}}}{1 \text{ a}}$$

$$= 3.15 \times 10^{7} \text{ s/a}$$

$$\text{Years in one mole of seconds} = \frac{6.02 \times 10^{23} \cancel{\text{s}}}{1 \text{ mol}} \times \frac{1 \text{ a}}{3.15 \times 10^{7} \cancel{\text{s}}}$$

$$= 1.91 \times 10^{16} \text{ a/mol}$$

It takes 1.91×10^{16} years to span a mole of seconds.

PRACTICE

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Understanding Concepts

9. (a) The molar mass of $H_2O_{(l)}$ is 18.01 g/mol.
   (b) The molar mass of $CO_{2(g)}$ is 44.01 g/mol.
   (c) The molar mass of $NaCl_{(s)}$ is 58.44 g/mol.
   (d) The molar mass of $C_{12}H_{22}O_{11(s)}$ is 342.34 g/mol.
   (e) The molar mass of $(NH_4)_2Cr_2O_{7(s)}$ is 252.10 g/mol.

   Note: Solutions for the rest of this course assume that students can correctly total a molar mass for a substance without showing work for the operation, unless specifically requested to do so. As a final example of the full work:

   $$1 \text{ mol } (NH_4)_2Cr_2O_7 = 2 \text{ mol N} + 8 \text{ mol H} + 2 \text{ mol Cr} + 7 \text{ mol O}$$

   $$M_{(NH_4)_2Cr_2O_7} = [(14.01 \times 2) + (1.01 \times 8) + (52.00 \times 2) + (16.00 \times 7)] \text{ g/mol}$$

   $$M_{(NH_4)_2Cr_2O_7} = 252.10 \text{ g/mol}$$

10. (a) $m_{C_{12}H_{22}O_{11}} = 10.00$ kg

   $M_{C_{12}H_{22}O_{11}} = 342.34$ g/mol

   $n_{C_{12}H_{22}O_{11}} = ?$

   $$n_{C_{12}H_{22}O_{11}} = 10.00 \text{ k}\cancel{\text{g}} \times \frac{1 \text{ mol}}{342.34 \cancel{\text{g}}}$$

   $$n_{C_{12}H_{22}O_{11}} = 0.02921 \text{ kmol} = 29.21 \text{ mol}$$

   The amount of sucrose is 29.21 mol.

   (b) $m_{NaCl} = 500$ g

   $M_{NaCl} = 58.44$ g/mol

   $n_{NaCl} = ?$

   $$n_{NaCl} = 500 \cancel{\text{g}} \times \frac{1 \text{ mol}}{58.44 \cancel{\text{g}}}$$

   $$n_{NaCl} = 8.56 \text{ mol}$$

   The amount of pickling salt (sodium chloride) is 8.56 mol.

   (c) $m_{C_3H_8} = 40.0$ g

   $M_{C_3H_8} = 44.11$ g/mol

   $n_{C_3H_8} = ?$

   $$n_{C_3H_8} = 40.0 \cancel{\text{g}} \times \frac{1 \text{ mol}}{44.11 \cancel{\text{g}}}$$

   $$n_{C_3H_8} = 0.907 \text{ mol}$$

The amount of propane is 0.907 mol (or 907 mmol).

(d) $m_{HC_9H_7O_4} = 325$ mg

$M_{HC_9H_7O_4} = 180.17$ g/mol

$n_{HC_9H_7O_4} = ?$

$$n_{HC_9H_7O_4} = 325 \text{ m}\cancel{\text{g}} \times \frac{1 \text{ mol}}{180.17 \cancel{\text{g}}}$$

$$n_{HC_9H_7O_4} = 1.80 \text{ mmol}$$

The amount of acetylsalicylic acid (Aspirin) is 1.80 mmol.

(e) Write $CH_3CHOHCH_3$ condensed to C_3H_8O for convenience.

$m_{C_3H_8O} = 150$ g

$M_{C_3H_8O} = 60.11$ g/mol

$n_{C_3H_8O} = ?$

$$n_{C_3H_8O} = 150 \cancel{\text{g}} \times \frac{1 \text{ mol}}{60.11 \cancel{\text{g}}}$$

$$n_{C_3H_8O} = 2.50 \text{ mol}$$

The amount of 2-propanol (rubbing alcohol) is 2.50 mol.

Note: The first printing of Chemistry 11 incorrectly gives the formula of 2-propanol as $CH_3CH_2OHCH_{3(l)}$. Using that formula would yield an answer of 2.45 mol.

11. (a) $n_{NH_3} = 4.22$ mol

$M_{NH_3} = 17.04$ g/mol

$m_{NH_3} = ?$

$$m_{NH_3} = 4.22 \cancel{\text{mol}} \times \frac{17.04 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{NH_3} = 71.9 \text{ g}$$

The mass of ammonia is 71.9 g.

(b) $n_{NaOH} = 0.224$ mol

$M_{NaOH} = 40.00$ g/mol

$m_{NaOH} = ?$

$$m_{NaOH} = 0.224 \cancel{\text{mol}} \times \frac{40.00 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{NaOH} = 8.96 \text{ g}$$

The mass of sodium hydroxide is 8.96 g.

(c) $n_{H_2O} = 57.3$ mmol

$M_{H_2O} = 18.02$ g/mol

$m_{H_2O} = ?$

$$m_{H_2O} = 57.3 \text{ m}\cancel{\text{mol}} \times \frac{18.02 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{H_2O} = 1.03 \times 10^3 \text{ mg} = 1.03 \text{ g}$$

The mass of water is 1.03 g.

(d) $n_{KMnO_4} = 9.44$ kmol

$M_{KMnO_4} = 158.04$ g/mol

$m_{KMnO_4} = ?$

$$m_{KMnO_4} = 9.44 \text{ k}\cancel{\text{mol}} \times \frac{158.04 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{KMnO_4} = 1.49 \times 10^3 \text{ kg} = 1.49 \text{ Mg}$$

The mass of potassium permanganate is 1.49 Mg (or 1.49 t).

(e) $n_{(NH_4)_2SO_4} = 0.77$ mol

$M_{(NH_4)_2SO_4} = 132.16$ g/mol

$m_{(NH_4)_2SO_4} = ?$

$$m_{(NH_4)_2SO_4} = 0.77 \cancel{\text{mol}} \times \frac{132.16 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{(NH_4)_2SO_4} = 1.0 \times 10^2 \text{ g} = 0.10 \text{ kg}$$

The mass of ammonium sulfate is 0.10 kg.

12. (a) $n_{CO_2} = 15$ mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$N_{CO_2} = ?$

$$N_{CO_2} = 15 \cancel{\text{mol}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \cancel{\text{mol}}}$$

$$N_{CO_2} = 9.0 \times 10^{24} \text{ molecules}$$

There are 9.0×10^{24} molecules of carbon dioxide in the sample.

(b) $m_{NH_3} = 15$ g

$M_{NH_3} = 17.04$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$N_{NH_3} = ?$

$$n_{NH_3} = 15 \cancel{\text{g}} \times \frac{1 \text{ mol}}{17.04 \cancel{\text{g}}}$$

$$n_{NH_3} = 0.88 \text{ mol}$$

$$N_{NH_3} = 0.88 \cancel{\text{mol}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \cancel{\text{mol}}}$$

$$N_{NH_3} = 5.3 \times 10^{23} \text{ molecules}$$

or

$$N_{NH_3} = 15 \cancel{\text{g}} \times \frac{1 \cancel{\text{mol}}}{17.04 \cancel{\text{g}}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \cancel{\text{mol}}}$$

$$N_{NH_3} = 5.3 \times 10^{23} \text{ molecules}$$

There are 5.3×10^{23} molecules of ammonia in the sample.

(c) $m_{HCl} = 15$ g

$M_{HCl} = 36.46$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$N_{HCl} = ?$

$$n_{HCl} = 15 \cancel{\text{g}} \times \frac{1 \text{ mol}}{36.46 \cancel{\text{g}}}$$

$$n_{HCl} = 0.41 \text{ mol}$$

$$N_{HCl} = 0.41 \cancel{\text{mol}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \cancel{\text{mol}}}$$

$$N_{HCl} = 2.5 \times 10^{23} \text{ molecules}$$

or

$$N_{HCl} = 15\ \cancel{g} \times \frac{1\ \cancel{mol}}{36.46\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}}$$

$$N_{HCl} = 2.5 \times 10^{23}\ \text{molecules}$$

There are 2.5×10^{23} molecules of hydrogen chloride in the sample.

(d) $m_{NaCl} = 15$ g

$M_{NaCl} = 58.44$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$N_{NaCl} = ?$

$$n_{NaCl} = 15\ \cancel{g} \times \frac{1\ \text{mol}}{58.44\ \cancel{g}}$$

$$n_{NaCl} = 0.26\ \text{mol}$$

$$N_{NaCl} = 0.26\ \cancel{mol} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}}$$

$$N_{NaCl} = 1.5 \times 10^{23}\ \text{formula units}$$

or

$$N_{NaCl} = 15\ \cancel{g} \times \frac{1\ \cancel{mol}}{58.44\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}}$$

$$N_{NaCl} = 1.5 \times 10^{23}\ \text{formula units}$$

There are 1.5×10^{23} formula units of sodium chloride in the sample.

13. (a) $M_{CO_2} = 44.01$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$m_{CO_2} = ?$

$$m_{CO_2} = \frac{44.01\ \text{g}}{1\ \cancel{mol}} \times \frac{1\ \cancel{mol}}{6.02 \times 10^{23}\ \text{molecules}}$$

$$m_{CO_2} = 7.31 \times 10^{-23}\ \text{g/molecule (on average)}$$

The average mass of a molecule in a sample of carbon dioxide from respiration is 7.31×10^{-23} g.

(b) $M_{C_6H_{12}O_6} = 180.18$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$m_{C_6H_{12}O_6} = ?$

$$m_{C_6H_{12}O_6} = \frac{180.18\ \text{g}}{1\ \cancel{mol}} \times \frac{1\ \cancel{mol}}{6.02 \times 10^{23}\ \text{molecules}}$$

$$m_{C_6H_{12}O_6} = 2.99 \times 10^{-22}\ \text{g/molecule (on average)}$$

The average mass of a molecule in a sample of glucose from photosynthesis is 2.99×10^{-22} g.

(c) $M_{O_2} = 32.00$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$m_{O_2} = ?$

$$m_{O_2} = \frac{32.00\ \text{g}}{1\ \cancel{mol}} \times \frac{1\ \cancel{mol}}{6.02 \times 10^{23}\ \text{molecules}}$$

$$m_{O_2} = 5.32 \times 10^{-23}\ \text{g/molecule (on average)}$$

The average mass of a molecule in a sample of oxygen from photosynthesis is 5.32×10^{-23} g.

14. m_{H_2O} = 1.000 $\cancel{L}$ × 1.00 kg/$\cancel{L}$ = 1.00 kg

M_{H_2O} = 18.02 g/mol

N_A = 6.02×10^{23} entities/mol

N_{H_2O} = ?

Note: It is assumed that students by this point will automatically apply the 1.00 g/mL, 1.00 kg/L, and 1.00 Mg/m^3 (1.00 t/m^3) conversions for the mass–volume relationship of pure water.

$$n_{H_2O} = 1.00\ \text{k}\cancel{\text{g}} \times \frac{1\ \text{mol}}{18.02\ \cancel{\text{g}}}$$

$$n_{H_2O} = 0.0555\ \text{kmol} = 55.5\ \text{mol}$$

$$N_{H_2O} = 55.5\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{H_2O} = 3.34 \times 10^{25}\ \text{molecules}$$

or

$$N_{H_2O} = 1.00\ \text{k}\cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{18.02\ \cancel{\text{g}}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{H_2O} = 3.34 \times 10^{25}\ \text{molecules}$$

There are 3.34×10^{25} molecules of water in 1.000 L.

SECTION 4.4 QUESTIONS

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Understanding Concepts

1. (a) n_{O_2} = 1.5 mol

N_A = 6.02×10^{23} entities/mol

N_{O_2} = ?

$$N_{O_2} = 1.5\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{O_2} = 9.0 \times 10^{23}\ \text{molecules}$$

There are 9.0×10^{23} molecules of oxygen in the sample.

(b) 2 atoms/molecule × 9.0×10^{23} molecules = 1.8×10^{24} atoms.

There are 1.8×10^{24} atoms of oxygen in the sample.

2. $m_{H_2C_6H_6O_6}$ = 90 mg

$M_{H_2C_6H_6O_6}$ = 176.14 g/mol

N_A = 6.02×10^{23} entities/mol

$N_{H_2C_6H_6O_6}$ = ?

$$n_{H_2C_6H_6O_6} = 90\ \text{m}\cancel{\text{g}} \times \frac{1\ \text{mol}}{176.14\ \cancel{\text{g}}}$$

$$n_{H_2C_6H_6O_6} = 0.51\ \text{mmol}$$

$$N_{H_2C_6H_6O_6} = 0.51\ \text{m}\cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{H_2C_6H_6O_6} = 3.1 \times 10^{20}\ \text{molecules}$$

or

$$N_{H_2C_6H_6O_6} = 90\ \text{m}\cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{176.14\ \cancel{\text{g}}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{H_2C_6H_6O_6} = 3.1 \times 10^{20}\ \text{molecules}$$

There are 3.1×10^{20} molecules of ascorbic acid in the tablet.

3. (a) $n_{Ca(IO_3)_2} = 1.00 \times 10^{-2}$ mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$N_I = ?$

$$N_{Ca(IO_3)_2} = 1.00 \times 10^{-2}\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{Ca(IO_3)_2} = 6.02 \times 10^{21}\ \text{formula units}$$

$$N_I = 6.02 \times 10^{21}\ \text{formula units} \times \frac{2\ \text{atoms}}{\text{formula unit}}$$

$$N_I = 1.20 \times 10^{22}\ \text{atoms}$$

There are 1.20×10^{22} atoms of iodine in the sample.

(b) $n_{Ca(IO_3)_2} = 1.00 \times 10^{-2}$ mol

$M_{Ca(IO_3)_2} = 389.88$ g/mol

$m_{Ca(IO_3)_2} = ?$

$$m_{Ca(IO_3)_2} = 1.00 \times 10^{-2}\ \cancel{\text{mol}} \times \frac{389.88\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$m_{Ca(IO_3)_2} = 3.90\ \text{g}$$

The mass of calcium iodate is 3.90 g.

4. $m_{H_2O} = 500$ g

$M_{H_2O} = 18.02$ g/mol

$n_{H_2O} = ?$

$$n_{H_2O} = 500\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{18.02\ \cancel{\text{g}}}$$

$$n_{H_2O} = 27.7\ \text{mol}$$

$m_{C_{12}H_{22}O_{11}} = 200$ g

$M_{C_{12}H_{22}O_{11}} = 342.34$ g/mol

$n_{C_{12}H_{22}O_{11}} = ?$

$$n_{C_{12}H_{22}O_{11}} = 200\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{342.34\ \cancel{\text{g}}}$$

$$n_{C_{12}H_{22}O_{11}} = 0.584\ \text{mol (or 584 mmol)}$$

$m_{HC_2H_3O_2} = 25$ g

$M_{HC_2H_3O_2} = 60.06$ g/mol

$n_{HC_2H_3O_2} = ?$

$$n_{HC_2H_3O_2} = 25\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{60.06\ \cancel{\text{g}}}$$

$$n_{HC_2H_3O_2} = 0.42\ \text{mol}$$

$m_{H_3C_6H_5O_7} = 15$ g

$M_{H_3C_6H_5O_7} = 192.14$ g/mol

$n_{H_3C_6H_5O_7} = ?$

$$n_{H_3C_6H_5O_7} = 15\ \cancel{g} \times \frac{1\ \text{mol}}{192.14\ \cancel{g}}$$

$$n_{H_3C_6H_5O_7} = 0.078\ \text{mol (or 78 mmol)}$$

m_{NaCl} = 5 g

M_{NaCl} = 58.44 g/mol

n_{NaCl} = ?

$$n_{NaCl} = 5\ \cancel{g} \times \frac{1\ \text{mol}}{58.44\ \cancel{g}}$$

$$n_{NaCl} = 0.09\ \text{mol}$$

Converted to amounts, the recipe is: 27.7 mol water, 0.584 mol sugar, 0.42 mol vinegar, 0.078 mol citric acid, and 0.09 mol salt.

5. (a) $v_{C_2H_5OH}$ = 17 mL

$d_{C_2H_5OH}$ = 0.789 g/mL (density)

$M_{C_2H_5OH}$ = 46.08 g/mol

N_A = 6.02×10^{23} entities/mol

$N_{C_2H_5OH}$ = ?

$$m_{C_2H_5OH} = 17\ \cancel{mL} \times \frac{0.789\ g}{1\ \cancel{mL}}$$

$$m_{C_2H_5OH} = 13\ g$$

$$n_{C_2H_5OH} = 13\ \cancel{g} \times \frac{1\ \text{mol}}{46.08\ \cancel{g}}$$

$$n_{C_2H_5OH} = 0.29\ \text{mol}$$

$$N_{C_2H_5OH} = 0.29\ \cancel{mol} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}}$$

$$N_{C_2H_5OH} = 1.8 \times 10^{23}\ \text{molecules}$$

or

$$N_{C_2H_5OH} = 17\ \cancel{mL} \times \frac{0.789\ \cancel{g}}{1\ \cancel{mL}} \times \frac{1\ \cancel{mol}}{46.08\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}}$$

$$N_{C_2H_5OH} = 1.8 \times 10^{23}\ \text{molecules}$$

There are 1.8×10^{23} molecules of ethanol in the beer.

(b) v_{Ni} = 0.72 cm^3

d_{Ni} = 8.90 g/cm^3 (density)

M_{Ni} = 58.69 g/mol

N_A = 6.02×10^{23} entities/mol

N_{Ni} = ?

$$m_{Ni} = 0.72\ \cancel{cm^3} \times \frac{8.90\ g}{\cancel{cm^3}}$$

$$m_{Ni} = 6.4\ g$$

$$n_{Ni} = 6.4\ \cancel{g} \times \frac{1\ \text{mol}}{58.69\ \cancel{g}}$$

$$n_{Ni} = 0.11\ \text{mol}$$

$$N_{Ni} = 0.11\ \cancel{mol} \times \frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \cancel{mol}}$$
$$N_{Ni} = 6.6 \times 10^{22}\ \text{atoms}$$

or

$$N_{Ni} = 0.72\ \cancel{cm^3} \times \frac{8.90\ \cancel{g}}{\cancel{cm^3}} \times \frac{1\ \cancel{mol}}{58.69\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \cancel{mol}}$$
$$N_{Ni} = 6.6 \times 10^{22}\ \text{atoms}$$

There are 6.6×10^{22} atoms of nickel in the quarter.

(c) $m_{H_2O} = 100\ \cancel{mL} \times 1.00\ g/\cancel{mL} = 100\ g$

$M_{H_2O} = 18.02$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$N_{H_2O} = ?$

$$n_{H_2O} = 100\ \cancel{g} \times \frac{1\ \text{mol}}{18.02\ \cancel{g}}$$
$$n_{H_2O} = 5.55\ \text{mol}$$
$$N_{H_2O} = 5.55\ \cancel{mol} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}}$$
$$N_{H_2O} = 3.34 \times 10^{24}\ \text{molecules}$$

or

$$N_{H_2O} = 100\ \cancel{g} \times \frac{1\ \cancel{mol}}{18.02\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}}$$
$$N_{H_2O} = 3.34 \times 10^{24}\ \text{molecules}$$

There are 3.34×10^{24} molecules of water in 100 mL.

Applying Inquiry Skills

6. **Experimental Design**

A sample of silver nitrate is dissolved and completely reacted with copper metal. The ratio of amounts of silver product and copper reactant are determined from mass measurements.

Materials

- silver nitrate
- copper wire
- pure water
- 250-mL beaker
- stirring rod
- centigram balance
- paper towels

Procedure

1. Measure the mass of a piece of coiled copper wire to 0.01 g.
2. Dissolve the silver nitrate in about 150 mL of water in the beaker.
3. Place the copper wire in the solution and let stand until the reaction is complete.
4. Remove the silver from the surface of the copper wire, wipe the wire dry, and measure its final mass to 0.01 g.
5. Rinse the silver crystals in the beaker with water, and drain (decant) as much water as possible.
6. Place the wet silver on a piece of paper towel to dry.
7. When dry, measure the mass of the silver to 0.01 g.
8. Dispose of materials as instructed.

Evidence

mass of copper (initial) _______ g
mass of copper (final) _______ g
mass of silver _______ g

Analysis

The masses of silver and copper are divided by their molar masses to convert the quantities from masses to amounts. The ratio, amount of copper: amount of silver, is then calculated.

Making Connections

7. The primary point in any report should be that the concept of the mole is essential to converting the easily measurable quantity of substances (mass) into numerical amounts that are the numerical quantities represented by formulas and equations. Actual numbers are far too large for convenience, so the mole is defined so as to make these conversions quick and easy. It is, in fact, not necessary to know what the value of a mole is, numerically, to do predictive and descriptive work in chemistry — any more than one needs to know how many salt grains are in a shaker.

4.5 PERCENTAGE COMPOSITION

PRACTICE

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Understanding Concepts

1. $m_C = 7.2\text{ g}$

   $m_H = 2.2\text{ g}$

   $m_O = 17.6\text{ g}$

   $m_{total} = 27.0\text{ g}$

$$\%\text{ C} = \frac{7.2\ \cancel{\text{g}}}{27.0\ \cancel{\text{g}}} \times 100\%$$

$$\%\text{ C} = 27\%$$

$$\%\text{ H} = \frac{2.2\ \cancel{\text{g}}}{27.0\ \cancel{\text{g}}} \times 100\%$$

$$\%\text{ H} = 8.1\%$$

$$\%\text{ O} = \frac{17.6\ \cancel{\text{g}}}{27.0\ \cancel{\text{g}}} \times 100\%$$

$$\%\text{ O} = 65.2\%$$

   The percentage composition of the compound is 27% carbon atoms, 8.1% hydrogen atoms, and 65.2% oxygen atoms by mass.

2. (a) $m_O = (30.80 - 8.40)\text{ g} = 22.40\text{ g}$

   (b) $m_C = 8.40\text{ g}$

   $m_O = 22.40\text{ g}$

   $m_{total} = 30.80\text{ g}\ \ (CO_{2(g)})$

$$\%\text{ C} = \frac{8.40\ \cancel{\text{g}}}{30.80\ \cancel{\text{g}}} \times 100\%$$

$$\%\text{ C} = 27.3\%$$

$$\%\text{ O} = \frac{22.40\ \cancel{\text{g}}}{30.80\ \cancel{\text{g}}} \times 100\%$$

$$\%\text{ O} = 72.7\%$$

   The percentage composition of carbon dioxide is 27.3% carbon atoms and 72.7% oxygen atoms by mass.

3. a) $m_1 = 3.12$ g

$m_{Cu} = 2.50$ g

$m_O = (3.12 – 2.50)$ g $= 0.62$ g

$$\% \text{ Cu} = \frac{2.50 \cancel{\text{g}}}{3.12 \cancel{\text{g}}} \times 100\%$$

$$\% \text{ Cu} = 80.1\%$$

$$\% \text{ O} = \frac{0.62 \cancel{\text{g}}}{3.12 \cancel{\text{g}}} \times 100\%$$

$$\% \text{ O} = 20\%$$

The percentage composition of compound 1 is 80.1% copper atoms and 20% oxygen atoms by mass.

$m_2 = 1.62$ g

$m_{Cu} = 1.44$ g

$m_O = (1.62 – 1.44)$ g $= 0.18$ g

$$\% \text{ Cu} = \frac{1.44 \cancel{\text{g}}}{1.62 \cancel{\text{g}}} \times 100\%$$

$$\% \text{ Cu} = 88.9\%$$

$$\% \text{ O} = \frac{0.18 \cancel{\text{g}}}{1.62 \cancel{\text{g}}} \times 100\%$$

$$\% \text{ O} = 11\%$$

The percentage composition of compound 2 is 88.9% copper atoms and 11% oxygen atoms by mass.

(b) The two compounds cannot be the same substance, because the proportions of the elements are different.

PRACTICE

(Page 182)

Making Connections

4. (a) The total mass is 9.7 kg. Mass units cancel in percentage calculations.

natural rubber is 1.8/9.7 $\times$100%	= 19%
carbon black is 2.3/9.7 $\times$100%	= 24%
steel cord is 0.5/9.7 $\times$100%	= 5%
polyester and nylon is 0.5/9.7 $\times$100%	= 5%
steel bead wire is 0.5/9.7 $\times$100%	= 5%
chemicals are 1.4/9.7 $\times$100%	= 14%
synthetic rubber is 2.7/9.7 $\times$100%	= 28%

(b) From a Goodyear web site, the synthetic:natural rubber ratios of:

Passenger car tires	55%:45%
Light truck tires	50%:50%
Racing car tires	65%:35%
Earthmover tires	20%:80%

(c) Synthetic rubber gives more flexibility and traction, while natural rubber gives durability. Racing car tires must have superior traction, for example, but often don't even last for one race.

(d) Charles Goodyear discovered the process of vulcanizing rubber, which made it stable at higher temperatures. This made the rubber tire for vehicles possible, which has shaped the entire evolution of technological society.

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Try This Activity : What Makes Popcorn Pop?

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Popcorn Initial Mass (g)	Final Mass (g)	Change in Mass (g)	Percentage Water (%)	Percentage of Popped Corn (%)
whole 18.16	16.30	1.86	10.2	~90
split crosswise 15.34	13.90	1.45	9.45	~20
split lengthwise 15.04	13.54	1.50	9.97	<1

(a) In general, the results seem to confirm that popcorn pops because the moisture inside the large part (endosperm) of the seed vaporizes, builds up the internal pressure, and suddenly breaks the hard outer coating (pericarp). The evidence of the popping percentage of the whole seeds versus the lengthwise-split seeds clearly illustrates this. The results for the crosswise-split seeds are inconclusive, perhaps because this splitting will, in some cases, include some of the endosperm and, in other cases, be restricted only to the bottom, germinating part of the seed.

PRACTICE

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Understanding Concepts

5. $m_H = 1.01 \text{ u} \times 2 = 2.02 \text{ u}$

$m_S = 32.06 \text{ u} \times 1 = 32.06 \text{ u}$

$m_O = 16.00 \text{ u} \times 4 = 64.00 \text{ u}$

$m_{H_2SO_{4(aq)}} = 98.08 \text{ u}$

$$\% \text{ H} = \frac{2.02 \cancel{\text{u}}}{98.08 \cancel{\text{u}}} \times 100\%$$

$$\% \text{ H} = 2.06\%$$

$$\% \text{ S} = \frac{32.06 \cancel{\text{u}}}{98.08 \cancel{\text{u}}} \times 100\%$$

$$\% \text{ S} = 32.69\%$$

$$\% \text{ O} = \frac{64.00 \cancel{\text{u}}}{98.08 \cancel{\text{u}}} \times 100\%$$

$$\% \text{ O} = 65.25\%$$

The percentage composition of $H_2SO_{4(aq)}$ is 2.06% hydrogen atoms, 32.69% sulfur atoms, and 65.25% oxygen atoms by mass.

6. $m_{Mg^{2+}} = 24.31 \text{ u} \times 1 = 24.31 \text{ u}$

$m_{Mg(OH)_{2(s)}} = 58.33 \text{ u}$

$$\% \text{ Mg}^{2+} = \frac{24.31 \cancel{\text{u}}}{58.33 \cancel{\text{u}}} \times 100\%$$

$$\% \text{ Mg}^{2+} = 41.68\%$$

The percentage, by mass, of magnesium ions in $Mg(OH)_{2(s)}$ is 41.68%.

7. $m_{Fe^{2+}} = 55.85 \text{ u} \times 1 = 55.85 \text{ u}$

$m_{O^{2-}} = 16.00 \text{ u} \times 1 = 16.00 \text{ u}$

$m_{FeO_{(s)}} = 71.85 \text{ u}$

$$\% \text{ Fe}^{2+} = \frac{55.85 \cancel{\text{u}}}{71.85 \cancel{\text{u}}} \times 100\%$$

$$\% \text{ Fe}^{2+} = 77.73\%$$

$$\% \; O^{2-} = \frac{16.00 \cancel{u}}{71.85 \cancel{u}} \times 100\%$$

$$\% \; O^{2-} = 22.27\%$$

$$m_{Fe^{3+}} = 55.85 \text{ u} \times 2 = 111.70 \text{ u}$$

$$m_{O^{2-}} = 16.00 \text{ u} \times 3 = 48.00 \text{ u}$$

$$m_{Fe_2O_{3(s)}} = 159.70 \text{ u}$$

$$\% \; Fe^{3+} = \frac{111.70 \cancel{u}}{159.70 \cancel{u}} \times 100\%$$

$$\% \; Fe^{3+} = 69.94\%$$

$$\% \; O^{2-} = \frac{48.00 \cancel{u}}{159.70 \cancel{u}} \times 100\%$$

$$\% \; O^{2-} = 30.06\%$$

The percentage composition of $FeO_{(s)}$ is 77.73% iron(II) ions, 22.27% oxide ions by mass. The percentage composition of $Fe_2O_{3(s)}$ is 69.94% iron(III) ions, 30.06% oxide ions by mass.

8. $m_N = 14.01 \text{ u} \times 3 = 42.03 \text{ u}$

$$m_{(NH_4)_3PO_4} = 149.12 \text{ u}$$

$$\% \; N = \frac{42.03 \cancel{u}}{149.12 \cancel{u}} \times 100\%$$

$$\% \; N = 28.19\%$$

The percentage, by mass, of nitrogen atoms in $(NH_4)_3PO_{4(s)}$ is 28.19%.

SECTION 4.5 QUESTIONS

(Page 184)

Understanding Concepts

1. The percentage composition of a new compound can be used to establish the correct formula.
2. The total mass of the compound sample is (33.5 + 30.4) g = 63.9 g.

$$\% \; K^+ = \frac{33.5 \cancel{g}}{63.9 \cancel{g}} \times 100\% = 52.4\%$$

$$\% \; Cl^- = \frac{30.4 \cancel{g}}{63.9 \cancel{g}} \times 100\% = 47.6\%$$

The percentage composition of the compound is 52.4% potassium ions and 47.6% chloride ions by mass.

3. The total mass of the compound sample is (23.0 + 16.0 + 32.0) g = 71.0 g.

$$\% \; Na^+ = \frac{23.0 \cancel{g}}{71.0 \cancel{g}} \times 100\% = 32.4\%$$

$$\% \; S = \frac{16.0 \cancel{g}}{71.0 \cancel{g}} \times 100\% = 22.5\%$$

$$\% \; O = \frac{32.0 \cancel{g}}{71.0 \cancel{g}} \times 100\% = 45.1\%$$

The percentage composition of the compound is 32.4% sodium ions, 22.5% sulfur atoms, and 45.1% oxygen atoms by mass.

4. $m_N = 14.01 \text{ u} \times 2 = 28.02 \text{ u}$

$$m_{NH_4NO_{3(s)}} = 80.06 \text{ u}$$

$$\% \; N = \frac{28.02 \cancel{u}}{80.06 \cancel{u}} \times 100\%$$

$\% \text{ N} = 35.00\%$

$m_{N} = 14.01 \text{ u} \times 2 = 28.02 \text{ u}$

$m_{(NH_4)_2SO_{4(s)}} = 132.16 \text{ u}$

$$\% \text{ N} = \frac{28.02 \cancel{\text{u}}}{132.16 \cancel{\text{u}}} \times 100\%$$

$\% \text{ N} = 21.20\%$

The percentage by mass of nitrogen atoms in ammonium nitrate is 35.00%. In ammonium sulfate this percentage is 21.20%.

5. $m_{2H_2O} = 2[(1.01 \text{ u} \times 2) + 16.00 \text{ u}] = 36.04 \text{ u}$

$m_{CaSO_4 \cdot 2H_2O_{(s)}} = 172.18 \text{ u}$

$$\% \text{ } H_2O = \frac{36.04 \cancel{\text{u}}}{172.18 \cancel{\text{u}}} \times 100\%$$

$\% \text{ } H_2O = 20.93\%$

$m_{H_2O} = [(1.01 \text{ u} \times 2) + 16.00 \text{ u}] = 18.02 \text{ u}$

$m_{CaSO_4 \cdot H_2O_{(s)}} = 154.16 \text{ u}$

$$\% \text{ } H_2O = \frac{18.02 \cancel{\text{u}}}{154.16 \cancel{\text{u}}} \times 100\%$$

$\% \text{ } H_2O = 11.69\%$

The percentage by mass of water molecules in calcium sulfate dihydrate is 20.93%. In calcium sulfate monohydrate this percentage is 11.69%.

Applying Inquiry Skills

6. (a) **Procedure**

1. Measure and record the mass of the empty crucible to 0.01 g.
2. Place the copper in the crucible and measure and record the total mass to 0.01 g.
3. Add excess sulfur to the crucible, and heat strongly in a fume hood until all the copper has reacted, and all the excess sulfur has burned off.
4. Allow the crucible and contents to cool, and measure and record the total mass to 0.01 g.

(b) **Analysis**

The masses of copper and of copper sulfide can be obtained by subtraction, and used to calculate the percent (by mass) of copper ions in the compound. Subtracting from 100 will give the mass percent of sulfide ions in this compound.

(c) **Evaluation**

The experimental design is judged to be adequate. It should easily provide dependable evidence from which to calculate the required answer to the question.

Making Connections

7. (a) Percentage by mass is routinely used for home products like medications (0.05% oxymetazoline hydrochloride in a nasal spray) or cleaners (3% sodium hypochlorite when packed, in bleach). Foodstuffs almost never have mass percent data. Instead they list mass of each component per “serving size.”

(b) Typical products that use a percentage other than mass include vinegar, which is typically 5 or 7% acetic acid by volume, and alcoholic beverages, which list alcohol (ethanol) as volume percent as well.

4.6 EMPIRICAL AND MOLECULAR FORMULAS

PRACTICE

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Understanding Concepts

1. Empirical means derived from observation and experimentation.
2. A molecular formula gives the number of each kind of atom or ion, as opposed to an empirical formula which gives the simplest numerical ratio of the component atoms and/or ions.
3. Different compounds can exist because the same number and kind of atoms are bonded together differently, like ethanol, CH_3CH_2OH, and dimethyl ether, CH_3OCH_3. These two different compounds have very different properties, but would have the same percentage composition.
4. Possible molecular formulas could be C_2H_6, C_3H_9, C_4H_{12}, or indeed, any compound with the general formula C_nH_{2n}, where n is any integer.
5. (a) NO_2
   (b) CO_2
   (c) CH_2O
   (d) C_3H_2Cl
6. Sodium chloride does not exist as molecules, but as a three-dimensional lattice of ions; so there is no such concept as a molecular formula for this, or any other, ionic compound. The same rule applies to network solid elements and compounds — the formulas we use are always the simplest ratio of component ions or atoms.

Try This Activity : Distinguish Between Empirical and Molecular Formulas

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Name	Empirical Formula	Molecular Formula
ethane	CH_3	C_2H_6
butane	C_2H_5	C_4H_{10}
hexane	C_3H_7	C_6H_{14}
ethene	CH_2	C_2H_4
butene	CH_2	C_4H_8
hexene	CH_2	C_6H_{12}

4.7 CALCULATING CHEMICAL FORMULAS

PRACTICE

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Understanding Concepts

1. Empirical formulas can be determined from mass percent information.
2. Molecular formulas can be determined if one also has molar mass information.
3. Assume a 100 g sample, for convenience.

$m_{K^+} = 28.9\text{ g} \qquad M_{K^+} = 39.10\text{ g/mol}$

$m_S = 23.7\text{ g} \qquad M_S = 32.06\text{ g/mol}$

$m_O = 47.4\text{ g} \qquad M_O = 16.00\text{ g/mol}$

$$n_{K^+} = 28.9\ \cancel{\text{g}} \times \frac{1\text{ mol}}{39.10\ \cancel{\text{g}}}$$

$$n_{K^+} = 0.739\text{ mol}$$

$$n_S = 23.7\ \cancel{g} \times \frac{1\ mol}{32.06\ \cancel{g}}$$

$$n_S = 0.739\ mol$$

$$n_O = 47.4\ \cancel{g} \times \frac{1\ mol}{16.00\ \cancel{g}}$$

$$n_O = 2.96\ mol$$

The mole ratio, $K^+ : S : O$, is 0.739 : 0.739 : 2.96.

Simplifying (dividing each value by the lowest), we obtain 1.00 : 1.00 : 4.01, or almost exactly 1 : 1 : 4, making the empirical formula $KSO_{4(s)}$.

4. (a) Assume a 100.0 g sample, for convenience.

$m_C = 40.87$ g $\qquad M_C = 12.01$ g/mol

$m_H = 3.72$ g $\qquad M_H = 1.01$ g/mol

$m_N = 8.67$ g $\qquad M_N = 14.01$ g/mol

$m_O = 24.77$ g $\qquad M_O = 16.00$ g/mol

$m_{Cl} = 21.98$ g $\qquad M_{Cl} = 35.45$ g/mol

$$n_C = 40.87\ \cancel{g} \times \frac{1\ mol}{12.01\ \cancel{g}}$$

$$n_C = 3.403\ mol$$

$$n_H = 3.72\ \cancel{g} \times \frac{1\ mol}{1.01\ \cancel{g}}$$

$$n_H = 3.68\ mol$$

$$n_N = 8.67\ \cancel{g} \times \frac{1\ mol}{14.01\ \cancel{g}}$$

$$n_N = 0.619\ mol$$

$$n_O = 24.77\ \cancel{g} \times \frac{1\ mol}{16.00\ \cancel{g}}$$

$$n_O = 1.548\ mol$$

$$n_{Cl} = 21.98\ \cancel{g} \times \frac{1\ mol}{35.45\ \cancel{g}}$$

$$n_{Cl} = 0.6200\ mol$$

The mole ratio, C:H:N:O:Cl, is 3.403 : 3.68 : 0.619 : 1.548 : 0.6200.

Simplifying (dividing each by 0.619), we obtain 5.50 : 5.95 : 1.00 : 2.50 : 1.00. This is obviously not an integral ratio, but two more of the values become integers upon doubling, which gives 11.0 : 11.9 : 2.00 : 5.00 : 2.00, or 11: 12 : 2 : 5 : 2, making the empirical formula for chloromycetin $C_{11}H_{12}N_2O_5Cl_{2(s)}$.

(b) Assume a 100.0 g sample, for convenience.

$m_C = 41.86$ g $\qquad M_C = 12.01$ g/mol

$m_H = 4.65$ g $\qquad M_H = 1.01$ g/mol

$m_N = 16.28$ g $\qquad M_N = 14.01$ g/mol

$m_O = 18.60$ g $\qquad M_O = 16.00$ g/mol

$m_S = 18.60$ g $\qquad M_S = 32.06$ g/mol

$$n_C = 41.86\ \cancel{g} \times \frac{1\ \text{mol}}{12.01\ \cancel{g}}$$

$$n_C = 3.485\ \text{mol}$$

$$n_H = 4.65\ \cancel{g} \times \frac{1\ \text{mol}}{1.01\ \cancel{g}}$$

$$n_H = 4.60\ \text{mol}$$

$$n_N = 16.28\ \cancel{g} \times \frac{1\ \text{mol}}{14.01\ \cancel{g}}$$

$$n_N = 1.162\ \text{mol}$$

$$n_O = 18.60\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$

$$n_O = 1.163\ \text{mol}$$

$$n_S = 18.60\ \cancel{g} \times \frac{1\ \text{mol}}{32.06\ \cancel{g}}$$

$$n_S = 0.5802\ \text{mol}$$

The mole ratio, C:H:N:O:S, is 3.485 : 4.60 : 1.162 : 1.163 : 0.5802.

Simplifying, we obtain 6.007 : 7.93 : 2.003 : 2.005 : 1.000. This is obviously nearly an integral ratio, or 6 : 8 : 2 : 2 : 1, making the empirical formula for sulfanilamide $C_6H_8N_2O_2S_{(s)}$.

5. (a) Assume a 100 g sample, for convenience.

m_P = (100 – 43.6) g = 56.4 g $\quad$ M_P = 30.97 g/mol

m_O = 43.6 g $\quad$ M_O = 16.00 g/mol

$$n_P = 56.4\ \cancel{g} \times \frac{1\ \text{mol}}{30.97\ \cancel{g}}$$

$$n_P = 1.82\ \text{mol}$$

$$n_O = 43.6\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$

$$n_O = 2.73\ \text{mol}$$

The mole ratio, P : O is 1.82 : 2.73.

Simplifying (dividing each value by the lowest), we obtain 1.00 : 1.50, which, when doubled, is 2 : 3, making the empirical formula $P_2O_{3(s)}$.

(b) Assume a 100 g sample, for convenience.

m_P = (100 – 56.6) g = 43.4 g $\quad$ M_P = 30.97 g/mol

m_O = 56.6 g $\quad$ M_O = 16.00 g/mol

$$n_P = 43.4\ \cancel{g} \times \frac{1\ \text{mol}}{30.97\ \cancel{g}}$$

$$n_P = 1.40\ \text{mol}$$

$$n_O = 56.6\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$

$$n_O = 3.54\ \text{mol}$$

The mole ratio, P : O, is 1.40 : 3.54.

Simplifying (dividing each value by the lowest), we obtain 1.00 : 2.53, which, when doubled, is 2.00 : 5.06, or 2:5, making the empirical formula $P_2O_{5(s)}$.

6. Assume a 100.0 g sample, for convenience.

m_C = (26.80 – 4.90) g = 21.90 g M_C = 12.01 g/mol

m_H = 4.90 g M_H = 1.01 g/mol

$$n_C = 21.90\ \cancel{g} \times \frac{1\ mol}{12.01\ \cancel{g}}$$

$$n_C = 1.823\ mol$$

$$n_H = 4.90\ \cancel{g} \times \frac{1\ mol}{1.01\ \cancel{g}}$$

$$n_H = 4.85\ mol$$

The mole ratio, C:H, is 1.823 : 4.85.
Simplifying, we obtain 1.000 : 2.66. This is obviously not an integral ratio, nor will doubling the ratio help. Multiplying the ratio through by three gives 3.000 : 7.98, or 3 : 8, making the empirical formula for propane $C_3H_{8(g)}$.

Applying Inquiry Skills

7. (a) **Experimental Design**

A sample of copper oxide will be reacted with carbon to burn away the oxygen and leave copper. The empirical formula of the oxide will be determined from the proportions of copper and oxygen in the oxide.

Procedure

1. Measure and record the mass of the empty crucible to 0.01 g.
2. Add the copper oxide sample to the crucible and measure and record the total mass to 0.01 g.
3. Add excess carbon to the crucible, and heat strongly until all the carbon has burned away, and only copper metal remains.
4. Allow the crucible and contents to cool, and measure and record the total mass to 0.01 g.

Evidence

mass of crucible	______ g
mass of crucible plus copper oxide	______ g
mass of crucible plus copper	______ g

Analysis

The masses of copper and oxygen in the sample are found by subtraction, if we assume the sample contained only copper and oxide ions. These masses are converted to moles, to allow determination of the mole ratio, and thus the formula.

(b) If the oxide is $Cu_2O_{(s)}$, then the evidence would show:

m_{Cu^+} = 63.55 u × 2 = 127.10 u

$m_{O^{2-}}$ = 16.00 u × 1 = 16.00 u

$m_{Cu_2O_{(s)}}$ = 143.10 u

$$\% \ Cu^+ = \frac{127.10\ \cancel{u}}{143.10\ \cancel{u}} \times 100\%$$

$$\% \ Cu^+ = 88.819\%$$

$$\% \ O^{2-} = \frac{16.00\ \cancel{u}}{143.10\ \cancel{u}} \times 100\%$$

$$\% \ O^{2-} = 11.18\%$$

The percentage composition of $Cu_2O_{(s)}$ is approximately 89% copper ions and 11% oxygen ions by mass. If the sample were this compound, the evidence would show these proportions.

PRACTICE

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Understanding Concepts

8. Assume one mole of compound, 58.1 g.

   mass % C = 62.0% M_C = 12.01 g/mol

   mass % H = 10.4% M_H = 1.01 g/mol

   mass % O = 27.5% M_O = 16.00 g/mol

   $$n_C = \frac{62.0}{100} \times 58.1\ \cancel{g} \times \frac{1\ \text{mol}}{12.01\ \cancel{g}}$$
   $$n_C = 3.00\ \text{mol}$$
   $$n_H = \frac{10.4}{100} \times 58.1\ \cancel{g} \times \frac{1\ \text{mol}}{1.01\ \cancel{g}}$$
   $$n_H = 5.98\ \text{mol}$$
   $$n_O = \frac{27.5}{100} \times 58.1\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$
   $$n_O = 0.999\ \text{mol}$$

   The integral mole ratio is 3 : 6 : 1, so the molecular formula is C_3H_6O.

9. Assume one mole of compound, 92.0 g.

   mass % N = 30.4% M_N = 14.01 g/mol

   mass % O = 69.6% M_O = 16.00 g/mol

   $$n_N = \frac{30.4}{100} \times 92.0\ \cancel{g} \times \frac{1\ \text{mol}}{14.01\ \cancel{g}}$$
   $$n_N = 2.00\ \text{mol}$$
   $$n_O = \frac{69.6}{100} \times 92.0\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$
   $$n_O = 4.00\ \text{mol}$$

   The integral mole ratio is 2 : 4, so the molecular formula is N_2O_4, and the name is dinitrogen tetraoxide.

Applying Inquiry Skills

10. (a) **Analysis**

   Assume one mole of compound, 180.2 g.

   mass % C = 40.0% M_C = 12.01 g/mol

   mass % H = 6.8% M_H = 1.01 g/mol

   mass % O = 53.2% M_O = 16.00 g/mol

   $$n_C = \frac{40.0}{100} \times 180.2\ \cancel{g} \times \frac{1\ \text{mol}}{12.01\ \cancel{g}}$$
   $$n_C = 6.00\ \text{mol}$$
   $$n_H = \frac{6.8}{100} \times 180.2\ \cancel{g} \times \frac{1\ \text{mol}}{1.01\ \cancel{g}}$$
   $$n_H = 12\ \text{mol}$$
   $$n_O = \frac{53.2}{100} \times 180.2\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$
   $$n_O = 5.99\ \text{mol}$$

   The integral mole ratio is 6 : 12 : 6, so the molecular formula of this carbohydrate is $C_6H_{12}O_6$.

PRACTICE

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Understanding Concepts

11. Natural source products usually contain traces of many other substances, and will normally vary somewhat in composition.

Making Connections

12. Vitamin D is produced naturally beneath the human skin surface by reaction of the body to exposure to sunlight. It is classed as an essential substance (a vitamin) because a lack of it causes a wide variety of deficiency diseases, the most notable of which is rickets. Rickets is characterized by a softening of the bones, causing limbs to grow bent, especially in children; and by a whole host of other symptoms. In fact, vitamin D is a steroid hormone that affects a great number of bodily functions.

    In 1921 Sir Edward Mellanby reported that dogs raised in the absence of sunlight could be kept healthy with a proper diet and that cod-liver oil contained some trace substance that prevented rickets. Cod-liver oil became a common tonic for people for the next several decades. Vitamin D (which actually has several forms) has since been synthesized, and is commonly available without prescription in tablet form in pharmacies. One of vitamin D's many aspects is that it aids bones to use calcium, so patients who take calcium supplements usually take vitamin D as well. Adding vitamin D to milk automatically ensures that the calcium in the milk is more useful and that small children receive an adequate amount of the vitamin itself.

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Explore an Issue   Debate: Are Natural Vitamins Better for Your Health?

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PRACTICE

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Understanding Concepts

13. A mass spectrometer can provide data that allows the determination of the molar mass of a substance.
14. Since CH would be 13.02 g/mol, and the molar mass is known to be 26 g/mol, the molecular formula must be C_2H_2(g).
15. Assume one mole of the compound ethane, 30.1 g.

    mass % C = 79.8% $M_C = 12.01$ g/mol

    mass % H = 20.2% $M_H = 1.01$ g/mol

    $$n_C = \frac{79.8}{100} \times 30.1\ \cancel{g} \times \frac{1\ \text{mol}}{12.01\ \cancel{g}}$$
    $$n_C = 2.00\ \text{mol}$$
    $$n_H = \frac{20.2}{100} \times 30.1\ \cancel{g} \times \frac{1\ \text{mol}}{1.01\ \cancel{g}}$$
    $$n_H = 6.02\ \text{mol}$$

    The integral mole ratio is 2 : 6, so the molecular formula of ethane from this evidence is $C_2H_{6(g)}$.

Applying Inquiry Skills

16. **Experimental Design**

    A sample of compound will be strongly heated to cause it to decompose to $CaO_{(s)}$ and $CO_{2(g)}$. The proportions of these two constituents will be determined from mass measurements.

Procedure

1. Measure and record the mass of the empty crucible to 0.01 g.
2. Add the compound sample to the crucible and measure and record the total mass to 0.01 g.
3. Heat strongly until all the sample has decomposed, and only $CaO_{(s)}$ remains.
4. Allow the crucible and contents to cool, and measure and record the total mass to 0.01 g.

Evidence

mass of crucible	______ g
mass of crucible plus compound	______ g
mass of crucible plus calcium oxide	______ g

Analysis

The masses of CaO and CO_2 in the sample are found by subtraction, if we assume the sample contained only these two constituents. These masses are converted to moles, to allow determination of the mole ratio, and thus the proportions.

Reflecting

17. The compound can be analyzed with a combustion analyzer and with a mass spectrometer to provide evidence to use in calculating its molecular formula.

SECTIONS 4.6 – 4.7 QUESTIONS

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Understanding Concepts

1. An empirical formula gives the simplest integral ratio of its constituent atoms, while a molecular formula also gives the correct actual number of atoms.
2. To calculate a molecular formula, the molar mass is needed.
3. To determine the empirical formula of a compound from the percentage composition, find the mass of each element in 100 g of the compound, using percentage composition; then find the amount in moles of each element by using the molar mass of the element; and then find the lowest integral ratio of atoms to determine the empirical formula.
4. The empirical formulas are, respectively:
   C_5H_4
   C_2H_6O
   NH_2
   O
   C_5H_7
   $C_6H_8N_2O_2S$
5. Ionic compounds do not exist as molecules, but as three-dimensional lattices of ions; so there is no such concept as a molecular formula for an ionic compound. The formulas we use are always the simplest ratio of component ions.
6. Assume a 100.00 g sample, for percentage conversion convenience.

$m_C = 40.00$ g $\qquad M_C = 12.01$ g/mol

$m_H = 6.71$ g $\qquad M_H = 1.01$ g/mol

$m_O = (100.00 - 40.00 - 6.71)$ g $= 53.29$ g $\qquad M_O = 16.00$ g/mol

$$n_C = 40.00\ \cancel{g} \times \frac{1\ \text{mol}}{12.01\ \cancel{g}}$$

$$n_C = 3.331\ \text{mol}$$

$$n_H = 6.71\ \cancel{g} \times \frac{1\ \text{mol}}{1.01\ \cancel{g}}$$

$$n_H = 6.64\ \text{mol}$$

$$n_O = 53.29\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$

$$n_O = 3.331\ \text{mol}$$

The mole ratio, C : H : O, is 3.331 : 6.64 : 3.331.

Simplifying (dividing each value by the lowest), we obtain a ratio of 1.000 : 1.99 : 1.000, or almost exactly 1 : 2 : 1, making the empirical formula for the lactic acid CH_2O.

(b) Given a molar mass for lactic acid of 90 g/mol, and a molar mass for CH_2O (the empirical formula) of 30.03 g/mol, the molecular formula must be triple the empirical ratio, or $C_3H_6O_3$.

Applying Inquiry Skills

7. (a) **Procedure**
   1. Use a centigram balance to measure the masses of a penny and of a quarter, to 0.01 g.
   2. Use a decigram balance to measure the total masses of the pennies and of the quarters, to 0.1 g.
   3. Divide the total mass of the pennies by the mass of one penny to calculate the number of pennies.
   4. Divide the total mass of the quarters by the mass of one quarter to calculate the number of quarters.
   5. Express the numbers of the two types of coins as a ratio.

   (b) A parallel procedure a scientist might use:
   1. Use a reference to find the molar masses of each kind of atom in the compound, to 0.01 g/mol.
   2. Use a combustion analyzer to measure the total masses of each kind of atom in the compound, to 0.01 g.
   3. Divide the total mass of each type of atom by the molar mass, to calculate the amount of atoms.
   4. Express the numbers of the different types of atoms as a simplest integral ratio.

CHAPTER 4 REVIEW

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Understanding Concepts

1. Every compound has a specific proportion of constituent substances. This led scientists to the concept of specific kinds of atoms for each element, and to comparing their combining masses to obtain a relative scale of masses of atoms.
2. The relative atomic mass would be (exactly) 24 u.
3. A hydrogen atom would be 18/12 of its current mass, or 1.5×1.01 u, or 1.52 u.
4. Relative atomic masses cannot be assigned correctly unless the combining ratio is correctly known; so the correct molecular formula is necessary.
5. Elements consist of atoms with identical numbers of protons and electrons. The nuclei of atoms of elements, however, may vary in numbers of neutrons, which has negligible effect on chemical properties, but does change the atom's mass significantly. For nearly every element there are several of these *isotopes*, and the average mass of atoms of such an element is a value that depends on the mass of these isotopes and also their proportion in nature. The classic example is chlorine, where roughly 3/4 of any sample will consist of chlorine-35 atoms (molar mass 35.00 g/mol), and roughly 1/4 of the sample will be chlorine-37 atoms (molar mass 37.00 g/mol). The molar mass of chlorine, then, is the *average* molar mass of all the chlorine atoms in a sample, which works out to 35.45 g/mol.
6. For silicon, with significant amounts of three isotopes, and if we assume the molar mass of isotopes is the same as the mass number, to two decimals (which is approximately valid, although the last digit is uncertain ...)

$$M = [(0.9221 \times 28.00) + (0.0470 \times 29.00) + (0.0309 \times 30.00)]\ \text{g/mol}$$

$$M = 28.11\ \text{g/mol} \qquad \text{(The actual value is 28.09 g/mol.)}$$

7. (a) 12 u exactly (by definition)
   (b) 12 g exactly (by definition)
   (c) Avogadro's constant is (by definition) the number of entities in exactly 12 g of pure carbon-12. This number is used to define the mole, which is the (numerical) amount of any substance that is this number of entities (atoms, ions, molecules, or formula units).
   (d) The symbol *M* represents the quantity, molar mass, in g/mol units.
8. (a) Formula: $CaCO_{3(s)}$

$$M = [(40.08) + (12.01) + (16.00 \times 3)]$$

$$M = 100.09\ \text{g/mol}$$

(b) Formula: $N_2O_{4(g)}$

$M = [(14.01 \times 2) + (16.00 \times 4)]$

$M = 92.02$ g/mol

(c) Formula: $Na_2CO_3{\cdot}10H_2O_{(s)}$

$M = [(22.99 \times 2) + (12.01) + (16.00 \times 3) + (18.02 \times 10)]$

$M = 286.19$ g/mol

Note: The last solution above assumes that students will have memorized the molar mass of water — an extremely useful quantity to commit to memory. For future study, memorization of the molar mass of carbon dioxide, 44.01 g/mol, will be almost as useful, because it is a product in so many reactions.

9. (a) $m_{NaCl} = 1.000$ kg

$M_{NaCl} = 58.44$ g/mol

$n_{NaCl} = 1.000\ \text{k}\cancel{\text{g}} \times \dfrac{1\ \text{mol}}{58.44\ \cancel{\text{g}}}$

$n_{NaCl} = 0.01711$ kmol = 17.11 mol

1.000 kg of table salt is 17.11 mol.

(b) $m_{CO_2} = 1.000$ kg

$M_{CO_2} = 44.01$ g/mol

$n_{CO_2} = 1.000\ \text{k}\cancel{\text{g}} \times \dfrac{1\ \text{mol}}{44.01\ \cancel{\text{g}}}$

$n_{CO_2} = 0.02272$ kmol = 22.72 mol

1.000 kg of dry ice is 22.72 mol.

(c) $m_{H_2O} = 1.000$ kg

$M_{H_2O} = 18.02$ g/mol

$n_{H_2O} = 1.000\ \text{k}\cancel{\text{g}} \times \dfrac{1\ \text{mol}}{18.02\ \cancel{\text{g}}}$

$n_{H_2O} = 0.05549$ kmol = 55.49 mol

1.000 kg of water is 55.49 mol.

10. (a) $n_{O_2} = 1.50$ mol

$M_{O_2} = 32.00$ g/mol

$m_{O_2} = 1.50\ \cancel{\text{mol}} \times \dfrac{32.00\ \text{g}}{1\ \cancel{\text{mol}}}$

$m_{O_2} = 48.0$ g

1.50 mol of liquid oxygen is 48.0 g.

(b) $n_{Hg} = 1.50$ mmol

$M_{Hg} = 200.59$ g/mol

$m_{Hg} = 1.50\ \text{m}\cancel{\text{mol}} \times \dfrac{200.59\ \text{g}}{1\ \cancel{\text{mol}}}$

$m_{Hg} = 301$ mg (or 0.301 g)

1.50 mmol of liquid mercury is 301 mg.

(c) $n_{Br_2} = 1.50$ kmol

$M_{Br_2} = 159.80$ g/mol

$m_{Br_2} = 1.50\ \text{k}\cancel{\text{mol}} \times \dfrac{159.80\ \text{g}}{1\ \cancel{\text{mol}}}$

$m_{Br_2} = 240$ kg

1.50 kmol of liquid bromine is 240 kg.

11.(a) $n_{HC_2H_3O_2} = 0.42$ mol

$N_A = 6.02 \times 10^{23}$ (entities)/mol

$$N_{HC_2H_3O_2} = 0.42\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}}{1\ \cancel{\text{mol}}}$$

$$N_{HC_2H_3O_2} = 2.5 \times 10^{23}$$

0.42 mol of acetic acid is 2.5×10^{23} molecules.

(b) $n_{CO} = 7.6 \times 10^{-4}$ mol

$N_A = 6.02 \times 10^{23}$ (entities)/mol

$$N_{CO} = 7.6 \times 10^{-4}\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}}{1\ \cancel{\text{mol}}}$$

$$N_{CO} = 4.6 \times 10^{20}$$

7.6×10^{-4} mol of carbon monoxide is 4.6×10^{20} molecules.

(c) $m_{CCl_4} = 100$ g

$M_{CCl_4} = 153.81$ g/mol

$N_A = 6.02 \times 10^{23}$ (entities)/mol

$$n_{CCl_4} = 100\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{153.81\ \cancel{\text{g}}}$$

$$n_{CCl_4} = 0.650\ \text{mol}$$

$$N_{CCl_4} = 0.650\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}}{1\ \cancel{\text{mol}}}$$

$$N_{CCl_4} = 3.91 \times 10^{23}$$

or

$$N_{CCl_4} = 100\ \cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{153.81\ \cancel{\text{g}}} \times \frac{6.02 \times 10^{23}}{1\ \cancel{\text{mol}}}$$

$$N_{CCl_4} = 3.91 \times 10^{23}$$

100 g of carbon tetrachloride is 3.91×10^{23} molecules.

(d) $m_{H_2S} = 100$ g

$M_{H_2S} = 34.08$ g/mol

$N_A = 6.02 \times 10^{23}$ (entities)/mol

$$n_{H_2S} = 100\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{34.08\ \cancel{\text{g}}}$$

$$n_{H_2S} = 2.93\ \text{mol}$$

$$N_{H_2S} = 2.93\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}}{1\ \cancel{\text{mol}}}$$

$$N_{H_2S} = 1.77 \times 10^{24}$$

or

$$N_{H_2S} = 100\ \cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{34.08\ \cancel{\text{g}}} \times \frac{6.02 \times 10^{23}}{1\ \cancel{\text{mol}}}$$

$$N_{H_2S} = 1.77 \times 10^{24}$$

100 g of hydrogen sulfide is 1.77×10^{24} molecules.

12. (a) $M_{C_{14}H_{18}N_2O_5} = [(12.01 \times 14) + (1.01 \times 18) + (14.01 \times 2) + (16.00 \times 5)]$ g/mol

$M_{C_{14}H_{18}N_2O_5} = 294.34$ g/mol

The molar mass of Aspartame is 294.34 g/mol

(b) $m_{C_{14}H_{18}N_2O_5} = 35$ mg

$M_{C_{14}H_{18}N_2O_5} = 294.34$ g/mol

$n_{C_{14}H_{18}N_2O_5} = 35 \text{ m}\cancel{\text{g}} \times \dfrac{1 \text{ mol}}{294.34 \ \cancel{\text{g}}}$

$n_{C_{14}H_{18}N_2O_5} = 0.12$ mmol

35 mg of Aspartame is 0.12 mmol.

(c) $n_{C_{14}H_{18}N_2O_5} = 0.12 \text{ mmol} = 1.2 \times 10^{-4}$ mol

$N_A = 6.02 \times 10^{23}$ (entities)/mol

$N_{C_{14}H_{18}N_2O_5} = 1.2 \times 10^{-4} \ \cancel{\text{mol}} \times \dfrac{6.02 \times 10^{23}}{1 \ \cancel{\text{mol}}}$

$N_{C_{14}H_{18}N_2O_5} = 7.2 \times 10^{19}$

$N_H = 7.2 \times 10^{19} \times 18$

$N_H = 1.3 \times 10^{21}$

35 mg of Aspartame contains 1.3×10^{21} hydrogen atoms.

13. (a) Assume one mole of compound, for convenience.

$m_{Na^+} = 22.99 \dfrac{\text{g}}{\cancel{\text{mol}}} \times 1 \ \cancel{\text{mol}} = 22.99$ g

$m_{NaN_{3(s)}} = 65.02$ g

$\% \ Na^+ = \dfrac{22.99 \ \cancel{\text{g}}}{65.02 \ \cancel{\text{g}}} \times 100\%$

$\% \ Na^+ = 35.36\%$

` The percentage, by mass, of sodium ions in $NaN_{3(s)}$ is 35.36%.

(b) Assume one mole of compound, for convenience.

$m_{Al^{3+}} = 26.98 \dfrac{\text{g}}{\cancel{\text{mol}}} \times 2 \ \cancel{\text{mol}} = 53.96$ g

$m_{Al_2O_{3(s)}} = 101.96$ g

$\% \ Al^{3+} = \dfrac{53.96 \ \cancel{\text{g}}}{101.96 \ \cancel{\text{g}}} \times 100\%$

$\% \ Al^{3+} = 52.92\%$

The percentage, by mass, of aluminum ions in $Al_2O_{3(s)}$ is 52.92%.

(c) Assume one mole of compound, for convenience.

$m_N = 14.01 \dfrac{\text{g}}{\cancel{\text{mol}}} \times 1 \ \cancel{\text{mol}} = 14.01$ g

$m_{C_8H_{11}NO_{2(s)}} = 153.20$ g

$\% \ N = \dfrac{14.01 \ \cancel{\text{g}}}{153.20 \ \cancel{\text{g}}} \times 100\%$

$\% \ N = 9.145\%$

The percentage, by mass, of nitrogen atoms in $C_8H_{11}NO_{2(aq)}$ is 9.145%.

14. (a) An empirical formula gives the simplest integral ratio of the constituent atoms of a compound, while a molecular formula also gives the correct actual number of atoms in one molecule.

(b) An empirical formula is just a proportion of atoms, and cannot give a correct number of atoms per molecule without evidence of the mass of the molecule — it is like saying Tom is twice as old as Harry, when, without the age of one, the other's age cannot be determined.

15. (a) Assume a 100.00 g sample, for percentage conversion convenience.

m_C = 49.5 g M_C = 12.01 g/mol

m_H = 5.15 g M_H = 1.01 g/mol

m_N = 28.9 g M_N = 14.01 g/mol

m_O = (by subtraction) = 16.5 g M_O = 16.00 g/mol

$$n_C = 49.5\ \cancel{g} \times \frac{1\text{ mol}}{12.01\ \cancel{g}}$$

$$n_C = 4.12\text{ mol}$$

$$n_H = 5.15\ \cancel{g} \times \frac{1\text{ mol}}{1.01\ \cancel{g}}$$

$$n_H = 5.10\text{ mol}$$

$$n_N = 28.9\ \cancel{g} \times \frac{1\text{ mol}}{14.01\ \cancel{g}}$$

$$n_N = 2.06\text{ mol}$$

$$n_O = 16.5\ \cancel{g} \times \frac{1\text{ mol}}{16.00\ \cancel{g}}$$

$$n_O = 1.03\text{ mol}$$

The mole ratio, C : H : N : O, is 4.12 : 5.10 : 2.06 : 1.03.

Simplifying (dividing each value by the lowest), we obtain a ratio of 4.00 : 4.95 : 2.00 : 1.00, or almost exactly 4 : 5 : 2 : 1, making the empirical formula for the caffeine $C_4H_5N_2O$.

(b) Since the molar mass of the caffeine is 195 g/mol, about double the value of 97.11g/mol that we get for $C_4H_5N_2O$, the molecular formula must be $C_8H_{10}N_4O_2$.

16. (a) m_C = 3.161 g M_C = 12.01 g/mol

m_H = 0.266 g M_H = 1.01 g/mol

m_O = 1.052 g M_O = 16.00 g/mol

$$n_C = 3.161\ \cancel{g} \times \frac{1\text{ mol}}{12.01\ \cancel{g}}$$

$$n_C = 0.2632\text{ mol}$$

$$n_H = 0.266\ \cancel{g} \times \frac{1\text{ mol}}{1.01\ \cancel{g}}$$

$$n_H = 0.263\text{ mol}$$

$$n_O = 1.052\ \cancel{g} \times \frac{1\text{ mol}}{16.00\ \cancel{g}}$$

$$n_O = 0.06575\text{ mol}$$

The mole ratio, C : H : O, is 0.2632 : 0.263 : 0.06575.

Simplifying (dividing each value by the lowest), we obtain a ratio of 4.00 : 4.00 : 1.00, making the empirical formula for the ester C_4H_4O.

(b) Since the molar mass of the ester is 136 g/mol, about double the value of 68.08 g/mol that we get for C_4H_4O, the molecular formula must be $C_8H_8O_2$.

17. Assume a 100.00 g sample, for percentage conversion convenience.

m_C = 63.2 g M_C = 12.01 g/mol

m_H = 5.26 g $\qquad M_H$ = 1.01 g/mol

m_O = 31.6 g $\qquad M_O$ = 16.00 g/mol

$$n_C = 63.2\ \cancel{g} \times \frac{1\ \text{mol}}{12.01\ \cancel{g}}$$

$$n_C = 5.26\ \text{mol}$$

$$n_H = 5.26\ \cancel{g} \times \frac{1\ \text{mol}}{1.01\ \cancel{g}}$$

$$n_H = 5.21\ \text{mol}$$

$$n_O = 31.6\ \cancel{g} \times \frac{1\ \text{mol}}{16.00\ \cancel{g}}$$

$$n_O = 1.98\ \text{mol}$$

Simplifying (dividing each value by the lowest), we obtain a ratio of 2.66 : 2.63 : 1.00, which is not an integral ratio. Tripling this ratio gives 7.97 : 7.89 : 3.00, or almost exactly 8 : 8 : 3, making the empirical formula for the vanillin $C_8H_8O_3$.

18. m_{Na^+} = 22.99 u × 3 = 68.97 u

m_P = 30.97 u × 1 = 30.97 u

m_O = 16.00 u × 4 = 64.00 u

$m_{Na_3PO_{4(aq)}}$ = 163.94 u

$$\%\ Na^+ = \frac{68.97\ \cancel{u}}{163.94\ \cancel{u}} \times 100\%$$

$$\%\ Na^+ = 42.07\%$$

$$\%\ P = \frac{30.97\ \cancel{u}}{163.94\ \cancel{u}} \times 100\%$$

$$\%\ P = 18.89\%$$

$$\%\ O = \frac{64.00\ \cancel{u}}{163.94\ \cancel{u}} \times 100\%$$

$$\%\ O = 39.04\%$$

The percentage composition of $Na_3PO_{4(aq)}$ is 42.07% sodium ions, 18.89% phosphorus atoms, and 39.04% oxygen atoms, by mass.

19. m_{Na^+} = 22.99 u × 3 = 68.97 u

m_{As} = 30.97 u × 1 = 74.92 u

m_O = 16.00 u × 4 = 64.00 u

$m_{Na_3AsO_{4(aq)}}$ = 207.89 u

$$\%\ Na^+ = \frac{68.97\ \cancel{u}}{207.89\ \cancel{u}} \times 100\%$$

$$\%\ Na^+ = 33.18\%$$

$$\%\ As = \frac{74.92\ \cancel{u}}{207.89\ \cancel{u}} \times 100\%$$

$$\%\ As = 36.04\%$$

$$\% \text{ O} = \frac{64.00 \cancel{u}}{207.89 \cancel{u}} \times 100\%$$

$$\% \text{ O} = 30.79\%$$

The percentage composition of $Na_3AsO_{4(aq)}$ is 33.18% sodium ions, 36.04% arsenic atoms, and 30.79% oxygen atoms, by mass.

Applying Inquiry Skills

20. (a) **Analysis**

m_{Zn} = (36.244 – 35.603) g = 0.641 g M_{Zn} = 65.38 g/mol

$m_{compound}$ = (36.933 – 35.603) g = 1.330 g

m_{Cl} = (1.330 – 0.641) g = 0.689 g M_{Cl} = 35.45 g/mol

$$n_{Zn^{2+}} = 0.641 \cancel{g} \times \frac{1 \text{ mol}}{65.38 \cancel{g}}$$

$$n_{Zn^{2+}} = 0.00980 \text{ mol}$$

$$n_{Cl^-} = 0.689 \cancel{g} \times \frac{1 \text{ mol}}{35.45 \cancel{g}}$$

$$n_{Cl^-} = 0.0194 \text{ mol}$$

Simplifying (dividing each value by the lowest), we obtain a ratio of 1.00 : 1.98, or almost exactly 1 : 2, making the empirical formula for the compound $ZnCl_2$.

(b) **Evaluation**

The design is simple and straightforward, and is likely to provide reliable evidence, from which the answer can be easily calculated.

21. (a) If the mass of the nail is measured before and after reaction, the mass difference will be copper, and can be used to determine the number of atoms.
(b) Besides the evidence from (a), the molar mass of copper and the value of Avogadro's constant will be needed.
(c) If the mass of copper(II) sulfate is initially measured, and the reaction is continued until all of the copper(II) ions react, then masses can be obtained to allow the determination of percentage by mass of copper in the compound.
22. (a) The crucible and lid are preheated to drive off any combustible or vaporizable material, so the mass of the crucible will not change later when heated.
(b) Handling hot equipment means wearing proper clothing, using heat-resistant gloves or tongs, and being careful not to set hot materials down on unprotected bench tops.
23. Subtracting the mass of the element from the mass of the oxide formed should give the mass of oxygen reacted.

Making Connections

24. Carbon monoxide and carbon dioxide are formed from the same elements; but one of these compounds is highly toxic to humans, and the other is not.
25. Moles are used for medical materials because the reaction quantities are the principal concern, and reaction equations are dependent on numerical amounts. For foods, masses are common because that is the easy and convenient way to measure how much food there is in a package or container.
26. Typical answers might include ...
(a) Food testing
Water treatment
Cement manufacturing
Plastics manufacturing
Prescription lens grinding
(b) Natural gas treatment labs
Water quality test labs
Steel composition analysis
Lubricant contamination analysis
Air quality test facilities

Exploring

27. Typical information ...
    Salicylate compounds were known to be painkillers by the 5th century B.C. In an attempt to find a relief from arthritic pain for his father, Felix Hoffmann first synthesized acetylsalicylic acid in 1893. The compound was named Aspirin and was being marketed by the Bayer corporation in 1897. By 1899 Heinrich Dreser was routinely using this drug to treat arthritis.
    GO TO www.science.nelson.com, Chemistry 11, Teacher Centre.

28. Typical information ...
    The most recent value for Avogadro's constant is obtained by precise measuring of atomic (ionic) sizes in metallic crystals, yielding a value with 9 significant digits — $6.022\ 141\ 99 \times 10^{23}$ entities.
    GO TO www.science.nelson.com, Chemistry 11, Teacher Centre.

29. Typical information ...
    Cheese has been made for at least 5500 years. The protein casein, in milk, clumps into curds upon the addition of rennin. Cottage cheese is sold in this initial stage. These curds may be heated, pressed, or strained to remove liquid whey. Cheddaring is the process of compressing curds to remove moisture. Many cheeses are also "ripened" by treatment with bacteria and/or moulds, which may be added into the cheese (internal treatment, like Roquefort) or rubbed on the surface (surface treatment, like Brie or Camembert). The texture and flavour of cheeses depends primarily on their moisture content and the agents used to ripen them. The holes in Swiss cheese are formed during ripening by gases produced by the bacteria used to create the characteristic flavour.
    GO TO www.science.nelson.com, Chemistry 11, Teacher Centre.