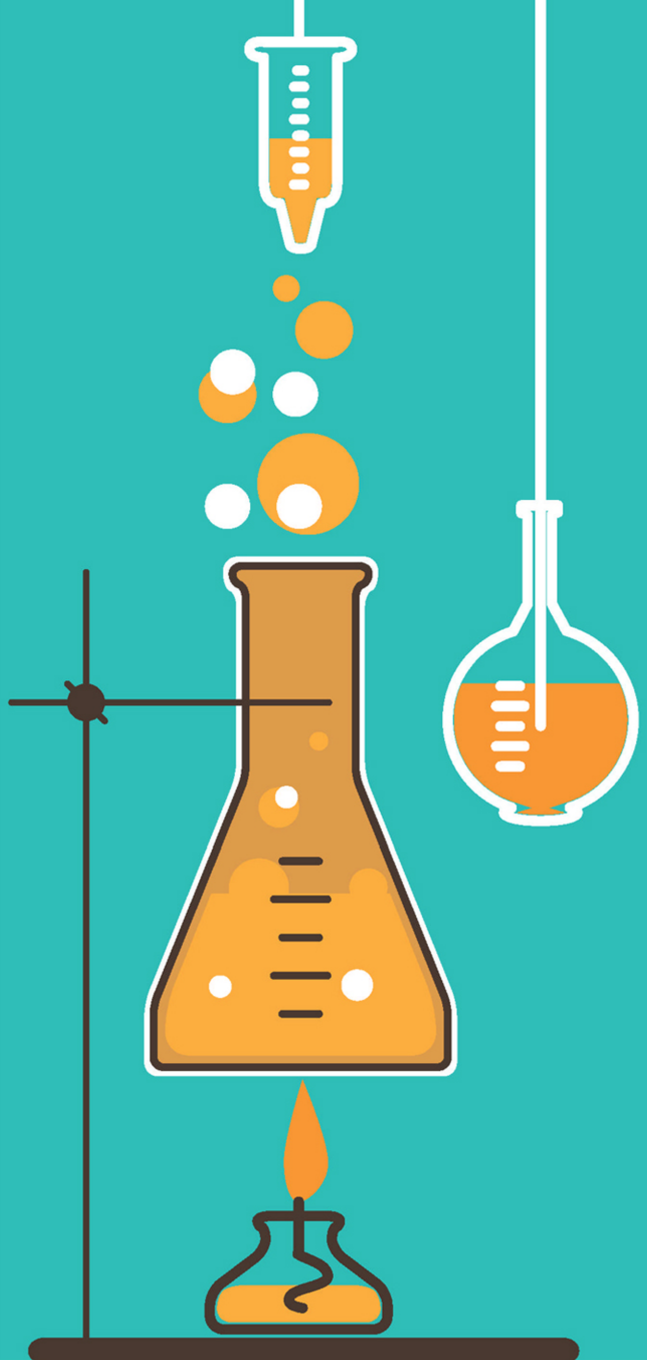

GLOBAL EDITION

Introductory Chemistry Essentials

SIXTH EDITION in SI Units

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Pearson

Periodic Table of the Elements

1	Atomic number
H	Element symbol
1.01	Atomic mass*
hydrogen	Element name

PERIOD	GROUP 1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B
1	1 **H** 1.01 hydrogen								
2	3 **Li** 6.94 lithium	4 **Be** 9.01 beryllium							
3	11 **Na** 22.99 sodium	12 **Mg** 24.31 magnesium							
4	19 **K** 39.10 potassium	20 **Ca** 40.08 calcium	21 **Sc** 44.96 scandium	22 **Ti** 47.88 titanium	23 **V** 50.94 vanadium	24 **Cr** 52.00 chromium	25 **Mn** 54.94 manganese	26 **Fe** 55.85 iron	27 **Co** 58.93 cobalt
5	37 **Rb** 85.47 rubidium	38 **Sr** 87.62 strontium	39 **Y** 88.91 yttrium	40 **Zr** 91.22 zirconium	41 **Nb** 92.91 niobium	42 **Mo** 95.95 molybdenum	43 **Tc** (99) technetium	44 **Ru** 101.07 ruthenium	45 **Rh** 102.91 rhodium
6	55 **Cs** 132.91 cesium	56 **Ba** 137.33 barium	57 **La** 138.91 lanthanum	72 **Hf** 178.49 hafnium	73 **Ta** 180.95 tantalum	74 **W** 183.85 tungsten	75 **Re** 186.21 rhenium	76 **Os** 190.23 osmium	77 **Ir** 192.22 iridium
7	87 **Fr** (223) francium	88 **Ra** (226) radium	89 **Ac** (227) actinium	104 **Rf** (261) rutherfordium	105 **Db** (262) dubnium	106 **Sg** (263) seaborgium	107 **Bh** (262) bohrium	108 **Hs** (265) hassium	109 **Mt** (266) meitnerium

Lanthanide series	58 **Ce** 140.12 cerium	59 **Pr** 140.91 praseodymium	60 **Nd** 144.24 neodymium	61 **Pm** (147) promethium	62 **Sm** 150.36 samarium	63 **Eu** 151.97 europium
Actinide series	90 **Th** (232) thorium	91 **Pa** (231) protactinium	92 **U** (238) uranium	93 **Np** (237) neptunium	94 **Pu** (244) plutonium	95 **Am** (243) americium

*The mass number of an important radioactive isotope—not the atomic mass—is shown in parentheses for those elements with no stable isotopes.

EXAMPLE 6.9 Mass Percent Composition

Calculate the mass percent of Cl in freon-114 ($C_2Cl_4F_2$).

SORT

You are given the molecular formula of freon-114 and asked to find the mass percent of Cl.

GIVEN: $C_2Cl_4F_2$

FIND: Mass % Cl

STRATEGIZE

You can use the information in the chemical formula to substitute into the mass percent equation and obtain the mass percent Cl.

SOLUTION MAP

$$\boxed{\text{chemical formula}} \longrightarrow \boxed{\text{mass \% Cl}}$$

$$\text{mass \% Cl} = \frac{4 \times \text{molar mass Cl}}{\text{molar mass } C_2Cl_4F_2} \times 100\%$$

RELATIONSHIPS USED

mass percent of element X =

$$\frac{\text{mass of element } X \text{ in 1 mol of compound}}{\text{mass of 1 mol of compound}} \times 100\%$$

(mass percent equation, introduced in this section)

SOLVE

Calculate the molar mass of freon-114 and substitute the values into the equation to find mass percent Cl.

SOLUTION

$$4 \times \text{molar mass Cl} = 4(35.45\text{ g}) = 141.8\text{ g}$$

$$\begin{aligned}\text{molar mass } C_2Cl_4F_2 &= 2(12.01) + 4(35.45) + 2(19.00)\\ &= 24.02 + 141.8 + 38.00\\ &= \frac{203.8\text{ g}}{\text{mol}}\end{aligned}$$

$$\begin{aligned}\text{mass \% Cl} &= \frac{4 \times \text{molar mass Cl}}{\text{molar mass } C_2Cl_4F_2} \times 100\%\\ &= \frac{141.8\text{ g}}{203.8\text{ g}} \times 100\%\\ &= 69.58\%\end{aligned}$$

CHECK

Check your answer. Are the units correct? Does the answer make physical sense?

The units (%) are correct. The answer makes physical sense. Mass percent composition should never exceed 100%. If your answer is greater than 100%, you have made an error.

▶ **SKILLBUILDER 6.9 | Mass Percent Composition**

Acetic acid ($HC_2H_3O_2$) is the active ingredient in vinegar. Calculate the mass percent composition of O in acetic acid.

▶ **FOR MORE PRACTICE** Example 6.20; Problems 79, 80, 81, 82, 85, 86.

CONCEPTUAL CHECKPOINT 6.7

Which compound has the highest mass percent of O? (You should not have to perform any detailed calculations to answer this question.)

(a) CrO

(b) CrO_2

(c) Cr_2O_3

CHEMISTRY AND HEALTH

Fluoridation of Drinking Water

In the early 1900s, scientists discovered that people whose drinking water naturally contained fluoride (F^-) ions had fewer cavities than people whose water did not. At appropriate levels, fluoride strengthens tooth enamel, which prevents tooth decay. In an effort to improve public health, fluoride has been artificially added to drinking water supplies since 1945. In the United States today, about 62% of the population drinks artificially fluoridated drinking water. The American Dental Association and public health agencies estimate that water fluoridation reduces tooth decay by 40 to 65%.

The fluoridation of public drinking water, however, is often controversial. Some opponents argue that fluoride is available from other sources—such as toothpaste, mouthwash, drops, and pills—and therefore should not be added to drinking water. Anyone who wants fluoride can get it from these optional sources, they argue, and the government should not impose fluoride on the population. Other opponents argue that the risks associated with fluoridation are too great. Indeed, too much fluoride can cause teeth to become brown and spotted, a condition known as dental fluorosis. Extremely high levels can lead to skeletal fluorosis, a condition in which the bones become brittle and arthritic.

The scientific consensus is that, like many minerals, fluoride shows some health benefits at certain levels—about 1–4 mg/day for adults—but can have detrimental effects at higher levels. Consequently, most major cities fluoridate their drinking water at a level of about 0.7 mg/L. Most adults drink between 1 and 2 L of water per day, so they receive beneficial amounts of fluoride from the water. Bottled water does not normally contain fluoride. Fluoridated bottled water can sometimes be found in the infant section of supermarkets.

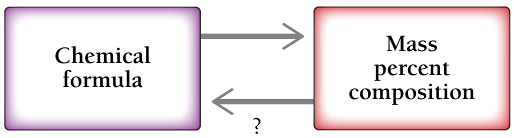
B6.1 CAN YOU ANSWER THIS? *Fluoride is often added to water as sodium fluoride (NaF). What is the mass percent composition of F^- in NaF? How many grams of NaF must be added to 1500 L of water to fluoridate it at a level of 0.7 mg F^-/L?*

6.8 Calculating Empirical Formulas for Compounds

- Determine an empirical formula from experimental data.
- Calculate an empirical formula from reaction data.

In Section 6.7, we learned how to calculate mass percent composition from a chemical formula. But can we go the other way? Can we calculate a chemical formula from mass percent composition? This is important because laboratory analyses of compounds do not often give chemical formulas directly; rather, they give the relative masses of each element present in a compound. For example, if we decompose water into hydrogen and oxygen in the laboratory, we could measure the masses of hydrogen and oxygen produced. Can we determine the chemical formula for water from this kind of data?

▶ We just learned how to go from the chemical formula of a compound to its mass percent composition. Can we also go the other way?



The answer is a qualified yes. We can determine a chemical formula, but it is the **empirical formula**, not the molecular formula. As we saw in Section 5.3, an empirical formula gives the smallest whole-number ratio of each type of atom in a compound, not the specific number of each type of atom in a molecule. Recall that the **molecular formula** is always a whole-number multiple of the empirical formula: Molecular formula = empirical $\times$ n, where $n = 1, 2, 3 \ldots$.

A chemical formula represents a ratio of atoms or moles of atoms, not a ratio of masses.

For example, the molecular formula for hydrogen peroxide is H_2O_2, and its empirical formula is HO.

$$\mathrm{HO} \times 2 \longrightarrow H_2O_2$$

Calculating an Empirical Formula from Experimental Data

Suppose we decompose a sample of water in the laboratory and find that it produces 3.0 g of hydrogen and 24 g of oxygen. How do we determine an empirical formula from these data?

▲ Water can be decomposed by an electric current into hydrogen and oxygen. How can we find the empirical formula for water from the masses of its component elements?

We know that an empirical formula represents a ratio of atoms or a ratio of moles of atoms, but it *does not* represent a ratio of masses. So the first thing we must do is convert our data from grams to moles. How many moles of each element formed during the decomposition? To convert to moles, we divide each mass by the molar mass of that element.

$$\text{mol H} = 3.0\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}} = 3.0\ \text{mol H}$$

$$\text{mol O} = 24\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 1.5\ \text{mol O}$$

From these data, we know there are 3 mol of H for every 1.5 mol of O. We can now write a pseudoformula for water:

$$H_3O_{1.5}$$

To get whole-number subscripts in our formula, we divide all the subscripts by the smallest one, in this case 1.5.

$$H_{\frac{3}{1.5}}O_{\frac{1.5}{1.5}} = H_2O$$

Our empirical formula for water, which in this case also happens to be the molecular formula, is H_2O. The following procedure can be used to obtain the empirical formula of any compound from experimental data. The left column outlines the procedure, and the center and right columns contain two examples of how to apply the procedure.

Obtaining an Empirical Formula from Experimental Data	EXAMPLE 6.10	EXAMPLE 6.11
	You decompose a compound containing nitrogen and oxygen in the laboratory and produce 24.5 g of nitrogen and 70.0 g of oxygen. Calculate the empirical formula of the compound.	A laboratory analysis of aspirin determines the following mass percent composition: C 60.00% H 4.48% O 35.53% Find the empirical formula.
1. Write down (or calculate) the masses of each element present in a sample of the compound. If you are given mass percent composition, assume a 100-g sample and calculate the masses of each element from the given percentages.	**GIVEN:** 24.5 g N 70.0 g O **FIND:** empirical formula	**GIVEN:** In a 100-g sample: 60.00 g C 4.48 g H 35.53 g O **FIND:** empirical formula
2. Convert each of the masses in Step 1 to moles by using the appropriate molar mass for each element as a conversion factor.	**SOLUTION** $24.5\ \cancel{\text{g N}} \times \frac{1\ \text{mol N}}{14.01\ \cancel{\text{g N}}} = 1.75\ \text{mol N}$ $70.0\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 4.38\ \text{mol O}$	**SOLUTION** $60.00\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 4.996\ \text{mol C}$ $4.48\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}} = 4.44\ \text{mol H}$ $35.53\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 2.221\ \text{mol O}$
3. Write down a pseudoformula for the compound, using the moles of each element (from Step 2) as subscripts.	$N_{1.75}O_{4.38}$	$C_{4.996}H_{4.44}O_{2.221}$
4. Divide all the subscripts in the formula by the smallest subscript.	$N_{\frac{1.75}{1.75}}O_{\frac{4.38}{1.75}} \longrightarrow N_1O_{2.5}$	$C_{\frac{4.996}{2.221}}H_{\frac{4.44}{2.221}}O_{\frac{2.221}{2.221}} \longrightarrow C_{2.25}H_2O_1$
5. If the subscripts are not whole numbers, multiply all the subscripts by a small whole number (see the following table) to arrive at whole-number subscripts.	$N_1O_{2.5} \times 2 \longrightarrow N_2O_5$ The correct empirical formula is N_2O_5.	$C_{2.25}H_2O_1 \times 4 \longrightarrow C_9H_8O_4$ The correct empirical formula is $C_9H_8O_4$.
	▶ **SKILLBUILDER 6.10** | A sample of a compound is decomposed in the laboratory and produces 165 g of carbon, 27.8 g of hydrogen, and 220.2 g O. Calculate the empirical formula of the compound.	▶ **SKILLBUILDER 6.11** | Ibuprofen, an aspirin substitute, has the mass percent composition: C 75.69%; H 8.80%; O 15.51%. Calculate the empirical formula of ibuprofen.
	▶ **FOR MORE PRACTICE** Problems 87, 88, 89, 90.	▶ **FOR MORE PRACTICE** Example 6.21; Problems 91, 92, 93, 94.

Fractional Subscript	Multiply by This Number to Get Whole-Number Subscripts
_.10	10
_.20	5
_.25	4
_.33	3
_.50	2
_.66	3
_.75	4

EXAMPLE 6.12 Calculating an Empirical Formula from Reaction Data

A 3.24-g sample of titanium reacts with oxygen to form 5.40 g of the metal oxide. What is the empirical formula of the metal oxide?

You are given the mass of titanium and the mass of the metal oxide that forms. You are asked to find the empirical formula. You need to recognize this problem as one requiring a special procedure and apply that procedure, which is outlined below.	**GIVEN:** 3.24 g Ti 5.40 g metal oxide **FIND:** empirical formula
1. Write down (or calculate) the masses of each element present in a sample of the compound. In this case, you are given the mass of the initial Ti sample and the mass of its oxide after the sample reacts with oxygen. The mass of oxygen is the difference between the mass of the oxide and the mass of titanium.	**SOLUTION** 3.24 g Ti mass O = mass oxide − mass titanium = 5.40 g − 3.24 g = 2.16 g O
2. Convert each of the masses in Step 1 to moles by using the appropriate molar mass for each element as a conversion factor.	$3.24\ \cancel{\text{g Ti}} \times \frac{1\text{ mol Ti}}{47.88\ \cancel{\text{g Ti}}} = 0.0677\text{ mol Ti}$ $2.16\ \cancel{\text{g O}} \times \frac{1\text{ mol O}}{16.00\ \cancel{\text{g O}}} = 0.135\text{ mol O}$
3. Write down a pseudoformula for the compound, using the moles of each element obtained in Step 2 as subscripts.	$Ti_{0.0677}O_{0.135}$
4. Divide all the subscripts in the formula by the smallest subscript.	$Ti_{\frac{0.0677}{0.0677}}O_{\frac{0.135}{0.0677}} \longrightarrow TiO_2$
5. If the subscripts are not whole numbers, multiply all the subscripts by a small whole number to arrive at whole-number subscripts.	As the subscripts are already whole numbers, this last step is unnecessary. The correct empirical formula is TiO_2.

▶ **SKILLBUILDER 6.12 | Calculating an Empirical Formula from Reaction Data**

A 1.56-g sample of copper reacts with oxygen to form 1.95 g of the metal oxide. What is the formula of the metal oxide?

▶ **FOR MORE PRACTICE** Problems 95, 96, 97, 98.

6.9 Calculating Molecular Formulas for Compounds

▶ Calculate a molecular formula from an empirical formula and molar mass.

We can determine the *molecular* formula of a compound from the empirical formula if we also know the molar mass of the compound. Recall from Section 6.8 that the molecular formula is always a whole-number multiple of the empirical formula.

$$\text{molecular formula} = \text{empirical formula} \times n, \text{ where } n = 1, 2, 3 \ldots$$

Suppose we want to find the molecular formula for fructose (a sugar found in fruit) from its empirical formula, CH_2O, and its molar mass, 180.2 g/mol. We know that the molecular formula is a whole-number multiple of CH_2O.

$$\text{molecular formula} = CH_2O \times n$$

▲ Fructose, a sugar found in fruit.

We also know that the molar mass is a whole-number multiple of the **empirical formula molar mass**, the sum of the masses of all the atoms in the empirical formula.

$$\text{molar mass} = \text{empirical formula molar mass} \times n$$

For a particular compound, the value of n in both cases is the same. Therefore, we can find n by calculating the ratio of the molar mass to the empirical formula molar mass.

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

For fructose, the empirical formula molar mass is:

$$\text{empirical formula molar mass} = 1(12.01) + 2(1.01) + 16.00 = 30.03\ \text{g/mol}$$

Therefore, n is:

$$n = \frac{180.2\ \text{g/mol}}{30.03\ \text{g/mol}} = 6$$

We can then use this value of n to find the molecular formula.

$$\text{molecular formula} = CH_2O \times 6 = C_6H_{12}O_6$$

EXAMPLE 6.13 Calculating Molecular Formula from Empirical Formula and Molar Mass

Naphthalene is a compound containing carbon and hydrogen that is used in mothballs. Its empirical formula is C_5H_4 and its molar mass is 128.16 g/mol. What is its molecular formula?

SORT

You are given the empirical formula and the molar mass of a compound and asked to find its molecular formula.

GIVEN: empirical formula $= C_5H_4$
molar mass $= 128.16$ g/mol

FIND: molecular formula

STRATEGIZE

In the first step, use the molar mass (which is given) and the empirical formula molar mass (which you can calculate based on the empirical formula) to determine n (the integer by which you must multiply the empirical formula to determine the molecular formula).

In the second step, multiply the subscripts in the empirical formula by n to arrive at the molecular formula.

SOLUTION MAP

molar mass, empirical formula molar mass → n

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

n, empirical formula → molecular formula

$$\text{molecular formula} = \text{empirical formula} \times n$$

SOLVE

First find the empirical formula molar mass. Next follow the solution map. Find n by dividing the molar mass by the empirical formula molar mass (which you just calculated). Multiply the empirical formula by n to determine the molecular formula.

SOLUTION

$$\text{empirical formula molar mass} = 5(12.01) + 4(1.01) = 64.09\ \text{g/mol}$$

$$n = \frac{\text{molar mass}}{\text{empirical formula mass}} = \frac{128.16\ \text{g/mol}}{64.09\ \text{g/mol}} = 2$$

$$\text{molecular formula} = C_5H_4 \times 2 = C_{10}H_8$$

CHECK

Check your answer. Does the answer make physical sense?

The answer makes physical sense because it is a whole-number multiple of the empirical formula. Any answer containing fractional subscripts would be an error.

▶ **SKILLBUILDER 6.13 | Calculating Molecular Formula from Empirical Formula and Molar Mass**

Butane is a compound containing carbon and hydrogen used as a fuel in butane lighters. Its empirical formula is C_2H_5, and its molar mass is 58.12 g/mol. Find its molecular formula.

▶ **SKILLBUILDER PLUS** A compound with the following mass percent composition has a molar mass of 60.10 g/mol. Find its molecular formula.

C 39.97% H 13.41% N 46.62%

▶ **FOR MORE PRACTICE** Example 6.22; Problems 99, 100, 101, 102.