Introductory Chemistry

Fourth Edition

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Chapter 6

Chemical Composition

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6.1 How Much Sodium?

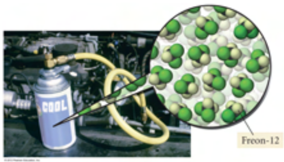
- Sodium is an important dietary mineral that we eat in our food, primarily as sodium chloride (table salt).
- Sodium is involved in the regulation of body fluids, and eating too much of it can lead to high blood pressure.

6.1 How Much Sodium?

- The FDA recommends a person consume less than 2.4 g (2400 mg) of ***sodium*** per day.
- The mass of ***sodium*** that we eat is not the same as the mass of ***sodium chloride*** that we eat.
- How many grams of sodium chloride can we consume and still stay below the FDA recommendation for sodium?
- The *chemical composition* of sodium chloride is given in its formula, NaCl.
- There is one sodium ion to every chloride ion.
- Since the masses of sodium and chlorine are different, the relationship between the mass of sodium and the mass of sodium chloride is not clear from the chemical formula alone.
- We need to calculate the amount of a constituent element in a given amount of a compound.

The information in a chemical formula, along with atomic and formula masses, can be used to calculate the amount of a constituent element in a compound.

How much iron is in a given amount of iron ore?
How much chlorine is in a given amount of a chlorofluorocarbon?



6.2 Counting by Weighing: Nails by the Pound

- Some hardware stores sell nails by the pound, which is easier than selling them by the nail.
- Customers often need hundreds of nails and counting them takes too long.
- A customer still needs to know the number of nails contained in a given weight of nails.
- This problem is similar to asking how many atoms are in a given mass of an element.

The hardware store customer buys 2.60 lb of medium-sized nails, where a dozen of these nails weigh 0.150 lb.
How many nails did the customer buy?

- The solution map for the problem is:

$$\boxed{\text{lb nails}} \xrightarrow{} \boxed{\text{doz nails}} \xrightarrow{} \boxed{\text{number of nails}}$$

$$\frac{1 \text{ doz nails}}{0.150 \text{ lb nails}} \qquad \frac{12 \text{ nails}}{1 \text{ doz nails}}$$

- We convert from lb to number of nails:

$$2.60\ \cancel{\text{lb nails}} \times \frac{1\ \cancel{\text{doz nails}}}{0.150\ \cancel{\text{lb nails}}} \times \frac{12 \text{ nails}}{1\ \cancel{\text{doz nails}}} = 208 \text{ nails}$$

6.2 Counting by Weighing: Nails by the Pound

- The conversion factor for the first part is the weight per dozen nails.

  0.150 lb nails = 1 doz nails
- The conversion factor for the second part is the number of nails in one dozen.

  1 doz nails = 12 nails

6.3 Counting by Weighing: Atoms by the Gram

- With atoms, we *must* use their mass as a way to count them.
- Atoms are too small and too numerous to count individually.
- Even if you could see atoms and counted them 24 hours a day as long as you lived, you would barely begin to count the number of atoms in something as small as a grain of sand.

6.3 Counting by Weighing: Atoms by the Gram

- With nails, we used a dozen as a convenient number in our conversions.
- A dozen is too small to use with atoms.
- We need a larger number because atoms are so small.
- The chemist's "dozen" is called the **mole (mol).**

$$1 \text{ mol} = 6.022 \times 10^{23}$$

One mole of anything is 6.022×10^{23} units of that thing.

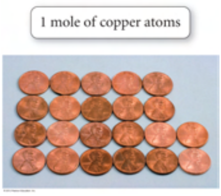
This number is called **Avogadro's number**, named after Amadeo Avogadro (1776–1856).

One mole of marbles corresponds to 6.022×10^{23} marbles.

One mole of sand grains corresponds to 6.022×10^{23} sand grains.

One mole of atoms, ions, or molecules generally makes up objects of reasonable size.

Twenty-two real *copper* pennies contain about 1 mol of copper (Cu) atoms.



Two large helium balloons contain approximately 1 mol of helium (He) atoms.

The size of the mole is an empirically measured quantity.

- *The numerical value of the mole is defined as being equal to the number of atoms in exactly 12 g of pure carbon-12.*
- This definition of the mole establishes a relationship between mass (grams of carbon) and number of atoms (Avogadro's number).
- This relationship allows us to count atoms by weighing them.

Converting moles to number of atoms: Convert 3.5 mol helium to the number of helium atoms.

GIVEN: 3.5 mol He

FIND: He atoms

RELATIONSHIPS USED 1 mol He = 6.022×10^{23} He atoms

SOLUTION MAP We draw a solution map showing the conversion from moles of He to He atoms.

$$\text{mol He} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{He atoms}}{1\ \text{mol He}}} \text{He atoms}$$

SOLUTION
Beginning with 3.5 mol He, we use the conversion factor to get to He atoms.

$$3.5\ \cancel{\text{mol He}} \times \frac{6.022 \times 10^{23}\ \text{He atoms}}{1\ \cancel{\text{mol He}}} = 2.1 \times 10^{24}\ \text{He atoms}$$

Converting number of atoms to moles:

Convert 1.1×10^{22} silver atoms to moles of silver.

GIVEN: 1.1×10^{22} Ag atoms

FIND: mol Ag

SOLUTION MAP

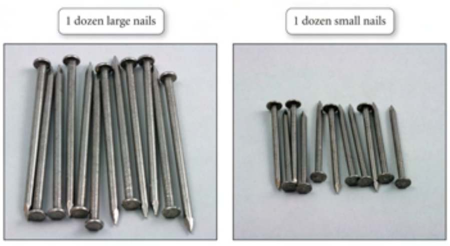
$$\text{Ag atoms} \xrightarrow{\frac{1\ \text{mol Ag}}{6.022 \times 10^{23}\ \text{Ag atoms}}} \text{mol Ag}$$

RELATIONSHIPS USED
1 mol Ag = 6.022×10^{23} Ag atoms (Avogadro's number)

SOLUTION

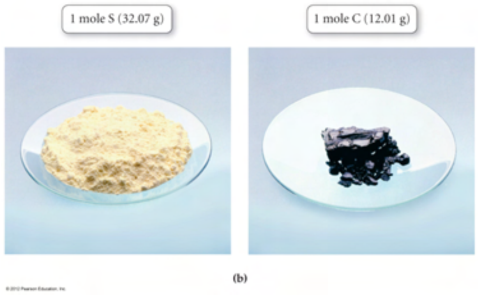
$$1.1 \times 10^{22}\ \cancel{\text{Ag atoms}} \times \frac{1\ \text{mol Ag}}{6.022 \times 10^{23}\ \cancel{\text{Ag atoms}}} = 1.8 \times 10^{-2}\ \text{mol Ag}$$

These pictures have the same number of nails. The weight of one dozen nails changes for different nails.



(a)

These pictures have the same number of atoms. The weight of one mole of atoms changes for different elements.



Molar Mass and Atomic Mass

- The atomic mass unit (amu) is defined as one-twelfth of the mass of a carbon-12 atom.
- The molar mass of any element—the mass of 1 mol of atoms of that element—is equal to the atomic mass of that element expressed in atomic mass units.
- One copper atom has an atomic mass of 63.55 amu.
- 1 mol of copper atoms has a mass of 63.55 g.
- The molar mass of copper is 63.55g/mol.

The mass of 1 mol of atoms of an element is its **molar mass.**

- The mass of 1 mol of atoms changes for different elements:

$$32.07 \text{ g sulfur} = 1 \text{ mol sulfur} = 6.022 \times 10^{23} \text{ S atoms}$$

$$12.01 \text{ g carbon} = 1 \text{ mol carbon} = 6.022 \times 10^{23} \text{ C atoms}$$

$$6.94 \text{ g lithium} = 1 \text{ mol lithium} = 6.022 \times 10^{23} \text{ Li atoms}$$

- The lighter the atom, the less mass in one mole of that atom.

Converting between Grams and Moles

Calculate the number of moles of sulfur in 57.8 g of sulfur.

GIVEN: 57.8 g S

FIND: mol S

SOLUTION MAP

$$\text{g S} \xrightarrow{\frac{1\ \text{mol S}}{32.07\ \text{g S}}} \text{mol S}$$

RELATIONSHIPS USED

32.07 g S = 1 mol S (molar mass of sulfur, from periodic table)

SOLUTION

$$57.8\ \cancel{\text{g S}} \times \frac{1\ \text{mol S}}{32.07\ \cancel{\text{g S}}} = 1.80\ \text{mol S}$$

Converting between Moles and Grams:

Calculate the mass of aluminum (g) in 6.73 moles of

GIVEN: 6.73 mol Al

FIND: g Al

SOLUTION MAP

$$\text{mol Al} \xrightarrow{\frac{26.98\ \text{g Al}}{1\ \text{mol Al}}} \text{g Al}$$

RELATIONSHIPS USED

26.98 g Al = 1 mol Al (molar mass of Al from periodic table)

SOLUTION

$$6.73\ \cancel{\text{mol Al}} \times \frac{26.98\ \text{g Al}}{1\ \cancel{\text{mol Al}}} = 182\ \text{g Al}$$

Converting between Grams and Number of Atoms:

How many aluminum atoms are in an aluminum can with a mass of 16.2 g?

GIVEN: 16.2 g Al

FIND: Al atoms

SOLUTION MAP

$$\text{g Al} \xrightarrow{\frac{1\ \text{mol Al}}{26.98\ \text{g Al}}} \text{mol Al} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{Al atoms}}{1\ \text{mol Al}}} \text{number of Al atoms}$$

RELATIONSHIPS USED

26.98 g Al = 1 mol Al (molar mass of aluminum, from periodic table)

6.022×10^{23} = 1 mol (Avogadro's number)

SOLUTION

$$16.2\ \cancel{\text{g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{26.98\ \cancel{\text{g Al}}} \times \frac{6.022 \times 10^{23}\ \text{Al atoms}}{1\ \cancel{\text{mol Al}}} = 3.62 \times 10^{23}\ \text{Al atoms}$$

6.4 Counting Molecules by the Gram

- For elements, the molar mass is the mass of 1 mol of atoms of that element.
- For compounds, the molar mass is the mass of 1 mol of molecules or formula units of that compound.
- Ionic compounds do not contain individual molecules.
- We convert between the mass of a compound and moles of the compound, then we calculate the number of molecules (or formula units) from moles.

Converting Between Grams and Moles of a Compound Requires the Molar Mass

- The molar mass of a compound in grams per mole is numerically equal to the formula mass of the compound in atomic mass units.
- The formula mass for a compound is the sum of the atomic masses of all of the atoms in a chemical formula.

The formula mass of CO_2 is found:

$$\begin{aligned}\text{Formula mass} &= 1(\text{Atomic mass of C}) + 2(\text{Atomic mass of O})\\ &= 1(12.01\ \text{amu}) + 2(16.00\ \text{amu})\\ &= 44.01\ \text{amu}\end{aligned}$$

The molar mass of CO_2 is:

$$\text{Molar mass} = 44.01\ \text{g/mol}$$

Converting Between Grams and Moles of a Compound: Calculate the mass in grams of 1.75 mol of water.

GIVEN: 1.75 mol H_2O

FIND: g H_2O

SOLUTION MAP

$$\text{mol } H_2O \xrightarrow{\frac{18.02\ \text{g } H_2O}{1\ \text{mol } H_2O}} \text{g } H_2O$$

Converting Between Grams and Moles of a Compound: Calculate the mass in grams of 1.75 mol of water.

RELATIONSHIPS USED

$$\begin{aligned} H_2O \text{ molar mass} &= 2(\text{Atomic mass H}) + 1(\text{Atomic mass O}) \\ &= 2(1.01) + 1(16.00) \\ &= 18.02\ \text{g/mol} \end{aligned}$$

SOLUTION

$$1.75\ \cancel{\text{mol } H_2O} \times \frac{18.02\ \text{g } H_2O}{\cancel{\text{mol } H_2O}} = 31.5\ \text{g } H_2O$$

Converting Between Number of Molecules and Mass of a Compound: What is the mass of 4.78×10^{24} NO_2 molecules?

GIVEN: 4.78×10^{24} NO_2 molecules

FIND: g NO_2

SOLUTION MAP

$$NO_2 \text{ molecules} \xrightarrow{\frac{1\ \text{mol } NO_2}{6.022 \times 10^{23}\ NO_2 \text{ molecules}}} \text{mol } NO_2 \xrightarrow{\frac{46.01\ \text{g } NO_2}{1\ \text{mol } NO_2}} \text{g } NO_2$$

Converting Between Number of Molecules and Mass of a Compound: What is the mass of 4.78×10^{24} NO_2 molecules?

RELATIONSHIPS USED

6.022×10^{23} molecules = 1 mol (Avogadro's number)

$$\begin{aligned} NO_2 \text{ molar mass} &= 1(\text{Atomic mass N}) + 2(\text{Atomic mass O}) \\ &= 14.01 + 2(16.00) \\ &= 46.01\ \text{g/mol} \end{aligned}$$

SOLUTION

$$4.78 \times 10^{24}\ \cancel{NO_2 \text{ molecules}} \times \frac{1\ \cancel{\text{mol } NO_2}}{6.022 \times 10^{23}\ \cancel{NO_2 \text{ molecules}}} \times \frac{46.1\ \text{g } NO_2}{1\ \cancel{\text{mol } NO_2}} = 365\ \text{g } NO_2$$

6.5 Chemical Formulas as Conversion Factors

3-Leaf Clover Analogy: How many leaves on 14 clovers?

3 leaves : 1 clover

3 leaves : 1 clover

clovers → leaves

$$\frac{3 \text{ leaves}}{1 \text{ clover}}$$

$$14 \cancel{\text{clovers}} \times \frac{3 \text{ leaves}}{1 \cancel{\text{clover}}} = 42 \text{ leaves}$$

6.5 Chemical Formulas as Conversion Factors

- The formula for carbon dioxide, CO_2, means there are two O atoms per CO_2 molecule.
- We write this as:

  2 O atoms : 1 CO_2 molecule
- Similarly,

  2 dozen O atoms : 1 dozen CO_2 molecules
- And:

  2 mol O: 1 mol CO_2

The conversion factor comes directly from the chemical formula.

8 legs : 1 spider

4 legs : 1 chair

2 H atoms : 1 H_2O molecule

Converting Between Moles of a Compound and Moles of a Constituent Element:
Find the number of moles of O in 1.7 mol $CaCO_3$.

GIVEN: 1.7 mol $CaCO_3$

FIND: mol O

SOLUTION MAP

$$\text{mol } CaCO_3 \xrightarrow{\frac{3 \text{ mol O}}{1 \text{ mol } CaCO_3}} \text{mol O}$$

RELATIONSHIPS USED

3 mol O : 1 mol $CaCO_3$ (from chemical formula)

SOLUTION

$$1.7\ \cancel{\text{mol } CaCO_3} \times \frac{3 \text{ mol O}}{1\ \cancel{\text{mol } CaCO_3}} = 5.1 \text{ mol O}$$

Converting Between Grams of a Compound and Grams of a Constituent Element: Find the mass of sodium in 15 g of NaCl.

- GIVEN: 15 g NaCl
- FIND: g Na
- SOLUTION MAP:

$$\text{g NaCl} \xrightarrow{\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}}} \text{mol NaCl} \xrightarrow{\frac{1 \text{ mol Na}}{1 \text{ mol NaCl}}} \text{mol Na} \xrightarrow{\frac{22.99 \text{ g Na}}{1 \text{ mol Na}}} \text{g Na}$$

- SOLUTION:

$$15\ \cancel{\text{g NaCl}} \times \frac{1\ \cancel{\text{mol NaCl}}}{58.44\ \cancel{\text{g NaCl}}} \times \frac{1\ \cancel{\text{mol Na}}}{1\ \cancel{\text{mol NaCl}}} \times \frac{22.99 \text{ g Na}}{1\ \cancel{\text{mol Na}}} = 5.9 \text{ g Na}$$

FIGURE 6.2 Mole relationships from a chemical formula

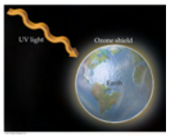
1 mol CCl_4 : 4 mol Cl

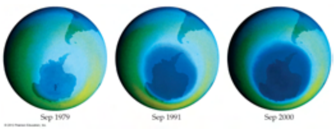
- The relationships inherent in a chemical formula allow us to convert between moles of the compound and moles of a constituent element (and vice versa).

Chemistry in the Environment
Chlorine in Chlorofluorocarbons

- Synthetic compounds known as chlorofluorocarbons (CFCs) are destroying a vital compound called ozone, O_3. in Earth's upper atmosphere.
- CFCs are chemically inert molecules used primarily as refrigerants and industrial solvents.
- In the upper atmosphere, sunlight breaks bonds within CFCs, resulting in the release of chlorine atoms.
- The chlorine atoms react with ozone and destroy it by converting it from O_3 into O_2.
- The thinning of ozone over populated areas is dangerous because ultraviolet light can harm living things and induce skin cancer in humans.
- Most developed nations banned the production of CFCs on January 1, 1996.
- CFCs still lurk in older refrigerators and air conditioning units and can leak into the atmosphere and destroy ozone.

- Upper atmospheric ozone is important because it acts as a shield to protect life on Earth from harmful ultraviolet light.
- Antarctic ozone levels in three Septembers from 1979 to 2000. The darkest blue colors indicate the lowest ozone levels.
- The mass percent chlorine changes from one type of chlorofluorocarbon to another.





6.6 Mass Percent Composition of Compounds

- The **mass percent composition,**

or **mass percent,** of an element

is the element's percentage of the total mass of

the compound.

$$\text{Mass percent of element } X = \frac{\text{Mass of } X \text{ in a sample of the compound}}{\text{Mass of the sample of the compound}} \times 100\%$$

Finding Mass Percent Composition

- A 0.358-g sample of chromium reacts with oxygen to form 0.523 g of the metal oxide.
- The mass percent of chromium is:

$$\text{Mass percent Cr} = \frac{\text{Mass Cr}}{\text{Mass metal oxide}} \times 100\%$$
$$= \frac{0.358\text{ g}}{0.523\text{ g}} \times 100\% = 68.5\%$$

Using Mass Percent Composition as a Conversion Factor

- We can use mass percent composition as a conversion factor between grams of a constituent element and grams of the compound.
- The mass percent composition of sodium in sodium chloride is 39%.

This can be written as:

39 g sodium : 100 g sodium chloride

Using Mass Percent Composition as a Conversion Factor

- The mass percent composition of sodium in sodium chloride is 39%.
- This can be written in fractional form:

$$\frac{39\text{ g Na}}{100\text{ g NaCl}} \quad \text{or} \quad \frac{100\text{ g NaCl}}{39\text{ g Na}}$$

These fractions are conversion factors between g Na and g NaCl.

The FDA recommends that adults consume less than 2.4 g of sodium per day.

How many grams of sodium chloride can you consume and still be within the FDA guidelines?

Sodium chloride is 39% sodium by mass.

GIVEN: 2.4 g Na

$\frac{39\text{ g Na}}{100\text{ g NaCl}}$

FIND: g NaCl

SOLUTION MAP

$\text{g Na} \xrightarrow{\frac{100\text{ g NaCl}}{39\text{ g Na}}} \text{g NaCl}$

RELATIONSHIPS USED

39 g Na : 100 g NaCl (given in the problem)

SOLUTION

$$2.4\ \cancel{\text{g Na}} \times \frac{100\text{ g NaCl}}{39\ \cancel{\text{g Na}}} = 6.2\text{ g NaCl}$$

6.7 Mass Percent Composition from a Chemical Formula

- Based on the chemical formula, the mass percent of element *X* in a compound is:

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

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

$$\text{Molar mass } C_2Cl_4F_2 = 2(12.01) + 4(35.45) + 2(19) = 24.02 + 141.8 + 38.00 = \frac{203.8\text{ g}}{\text{mol}}$$

$$\text{Mass \% Cl} = \frac{4 \times \text{Molar mass Cl}}{\text{Molar mass } C_2Cl_4F_2} \times 100\% = \frac{141.8\text{ g}}{203.8\text{ g}} \times 100\% = 69.58\%$$

Mass Percent Composition from a Chemical Formula:

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

GIVEN: $C_2Cl_4F_2$

FIND: Mass % Cl

SOLUTION MAP

$\text{Chemical formula} \rightarrow \text{mass \% Cl}$

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

Mass Percent Composition from a Chemical Formula:

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

RELATIONSHIPS USED

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

(mass percent equation, introduced in this section)

SOLUTION

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

$$\begin{aligned}\text{Molar mass } C_2Cl_4F_2 &= 2(12.01) + 4(35.45) + 2(19)\\ &= 24.02 + 141.8 + 38.00\\ &= \frac{203.8\text{ g}}{\text{mol}}\end{aligned}$$

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

Chemistry and Health
Fluoridation of Drinking Water

- Fluoride strengthens tooth enamel, which prevents tooth decay.
- 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.
- 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.
- Adults who drink between 1 and 2 L of water per day would receive the beneficial amounts of fluoride from the water.

Chemistry and Health
Fluoridation of Drinking Water

- *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 should be added to 1500 L of water to fluoridate it at a level of 1.0 mg F^-/L?*

6.8 Calculating Empirical Formulas for Compounds

- Can we calculate a chemical formula from mass percent composition?
- Yes, but it is the *empirical formula,* not the molecular formula.
- An empirical formula only 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.
- The molecular formula is always a whole-number multiple of the empirical formula.
- For example, the molecular formula for hydrogen peroxide is H_2O_2 and its empirical formula is HO.

Calculating an Empirical Formula from Experimental Data: Decomposition of Water

- 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?

Calculating an Empirical Formula from Experimental Data: Decomposition of Water to 3.0 g H and 24 g O.

- How many moles of each element formed during the decomposition of water?
- Divide the experimental mass of each element by the molar mass of that element.

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

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

- There are 3 mol of H for every 1.5 mol of O.

Calculating an Empirical Formula from Experimental Data: Decomposition of Water to 3.0 g H and 24 g O.

- Write a pseudo-formula for water:

$$H_3O_{1.5}$$

- To get whole-number subscripts in our formula, 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.

Obtaining an Empirical Formula from Experimental Data

1. Write down (or calculate) as given 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.
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. Write down a pseudo-formula for the compound, using the moles of each element (from Step 2) as subscripts.
4. Divide all the subscripts in the formula by the smallest subscript.
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.

If, after dividing by the smallest number of moles, the subscripts are not whole numbers, multiply all the subscripts by a small whole number to arrive at whole-number subscripts.

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

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?

GIVEN: 3.24 g Ti
5.40 g metal oxide
FIND: empirical formula

You cannot convert mass of metal oxide into moles because you would need its formula, and that is what you are trying to find.

You are given the mass of the initial Ti sample and the mass of its oxide after the sample reacts with oxygen. The difference is the mass of oxygen that combined with the titanium.

Calculating an Empirical Formula from Reaction Data

- To find the mass of oxygen, subtract the mass of titanium from the mass of the "metal oxide."

  The difference is the mass of oxygen.

  Mass Ti = 3.24 g Ti
  Mass O = Mass oxide – Mass titanium
  = 5.40 g Ti & O – 3.24 g Ti
  = 2.16 g O

- Now you can convert the mass of each element to moles.

6.9 Calculating Molecular Formulas for Compounds

- The molecular formula is always a whole-number multiple of the empirical formula.
- We need to find *n* in the expression:

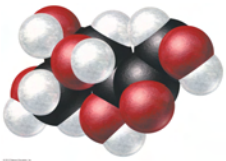
  Molecular formula = $CH_2O \times n$
- We can find *n* in the expression:

  Molar mass = Empirical formula molar mass × *n*
- Solving for n

$$n = \frac{\text{Molar mass}}{\text{Empirical formula molar mass}}$$

6.9 Calculating Molecular Formulas for Compounds: Fructose

- 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.



- The molecular formula is a whole-number multiple of CH_2O.

6.9 Calculating Molecular Formulas for Compounds: Fructose

- For fructose, the empirical formula molar

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

- Therefore, n is: $n = \dfrac{180.2\ \text{g/mol}}{30.03\ \text{g/mol}} = 6$

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

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

6.9 Calculating Molecular Formulas for Compounds

- 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 get the molecular formula).
- Multiply the subscripts in the empirical formula by n to arrive at the molecular formula.

Chapter 6 in Review

The mole concept:

- The mole is a specific number (6.022 x 10^{23}) that allows us to easily count atoms or molecules by weighing them.
- One mole of any element has a mass equivalent to its atomic mass in grams.
- One mole of any compound has a mass equivalent to its formula mass in grams.
- The mass of 1 mol of an element or compound is its molar mass.

Chapter 6 in Review

Chemical formulas and chemical composition:

- Chemical formulas indicate the relative number of each kind of element in a compound.
- These numbers are based on atoms or moles.
- By using molar masses, the information in a chemical formula can be used to determine the relative masses of each kind of element in a compound.
- The total mass of a sample of a compound can be related to the masses of the constituent elements contained in the compound.

Chapter 6 in Review

Empirical and molecular formulas from laboratory data:

- We can refer to the relative masses of each kind of element within a compound to determine the empirical formula of the compound.
- If the chemist also knows the molar mass of the compound, he or she can also determine its molecular formula.

Chemical Skills

- Converting between moles and number of atoms
- Converting between grams and moles
- Converting between grams and number of atoms or molecules
- Converting between moles of a compound and moles of a constituent element
- Converting between grams of a compound and grams of a constituent element
- Using mass percent composition as a conversion factor
- Determining mass percent composition from a chemical formula
- Determining an empirical formula from experimental data
- Calculating a molecular formula from an empirical formula and molar mass