

Chapter 6

Chemical Composition

Based on slides provided with Introductory Chemistry, Fifth Edition
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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.



How Much Sodium in Sodium Chloride?

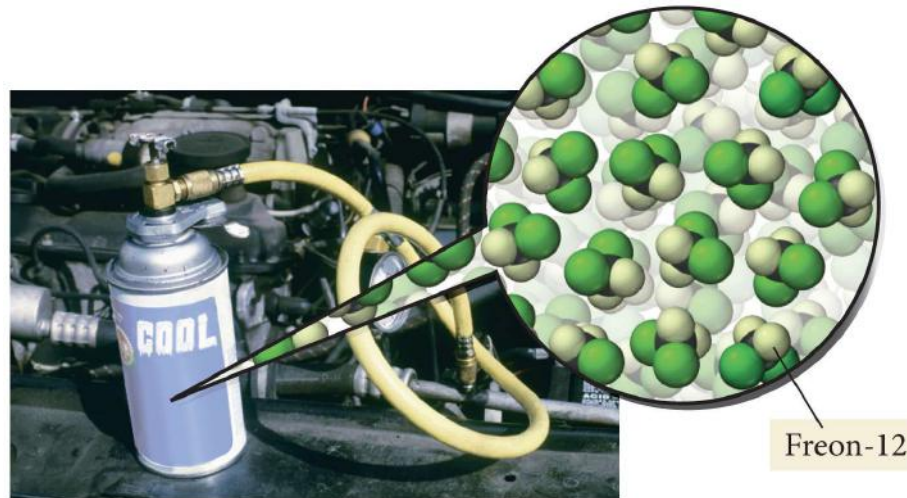
- The FDA recommends a person consume less than 2.4 g (2400 mg) of **sodium** per day.
- The mass of **sodium** that we eat is not the same as the mass of **sodium chloride** that we eat.
- How many grams of sodium chloride can we consume and still stay below the FDA recommendation for sodium?
- The *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?



Counting by Weighing: Nails by the Pound

Some hardware stores sell nails by the pound, which is easier than selling them by the nail.



3.4 lb nails

This problem is similar to asking how many atoms are in a given mass of an element.



8.25 grams carbon

A customer buys 2.60 lb of medium-sized nails, and a dozen of these nails weigh 0.150 lb. How many nails did the customer buy?

- The solution map for the problem is as follows:



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

- We convert from pounds to number of nails:

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

Counting by Weighing: Nails by the Pound

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

$$0.150 \text{ lb nails} = 1 \text{ doz nails}$$

- The conversion factor for the second part is the number of nails in one dozen.

$$1 \text{ doz nails} = 12 \text{ nails}$$

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.

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)**.

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

Avogadro's Number

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.

1 mole of copper atoms



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

1 mole of helium atoms



The Size of the Mole is a 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 also means: $1 \text{ g} = 6.022 \times 10^{23}$ atomic mass units*
- This definition of the mole establishes a relationship between mass (grams of carbon) and number of atoms (via 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.

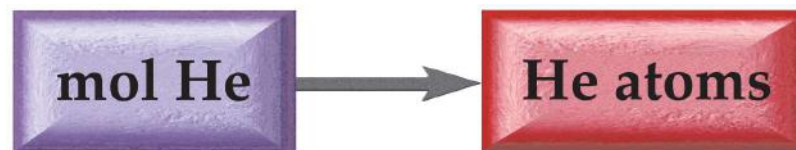
GIVEN: 3.5 mol He

FIND: He atoms

RELATIONSHIPS USED

1 mol He = 6.022×10^{23} He atoms

SOLUTION MAP



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

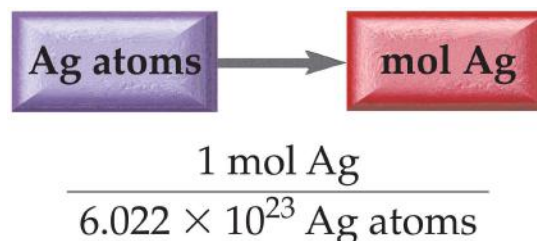
$$3.5 \text{ mol He} \times \frac{6.022 \times 10^{23} \text{ He atoms}}{1 \text{ mol He}} = 2.1 \times 10^{24} \text{ He atoms}$$

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



RELATIONSHIPS USED

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

SOLUTION

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

Practice:

Check your solution on next page

Calculate the number of atoms in 2.45 mol of copper

Practice: Calculate the number of atoms in 2.45 mol of copper

Given: Find:	2.45 mol Cu atoms Cu
Conceptual Plan:	$\text{mol Cu} \xrightarrow{\frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}}} \text{atoms Cu}$
Relationships:	1 mol = 6.022×10^{23} atoms
Solution:	$2.45 \cancel{\text{ mol Cu}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \cancel{\text{ mol}}}$ $= 1.48 \times 10^{24} \text{ atoms Cu}$
Check:	because atoms are small, the large number of atoms makes sense

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

1 dozen large nails



1 dozen small nails



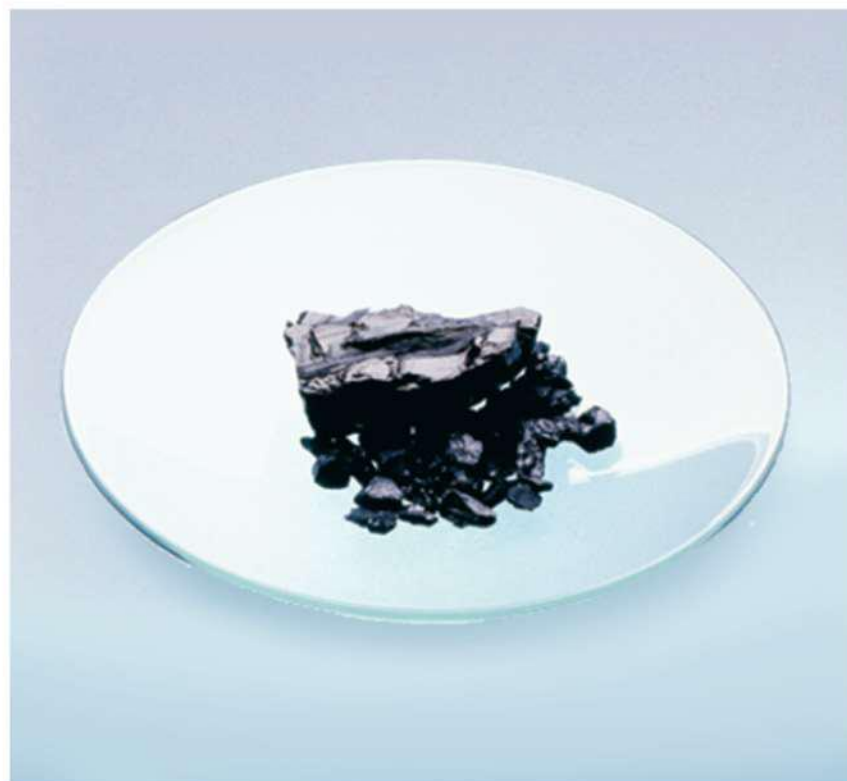
(a)

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

1 mole S (32.07 g)



1 mole C (12.01 g)



(b)

Molar Mass and Atomic Mass

- The atomic mass unit (amu) is defined as one-twelfth of the mass of a carbon-12 atom.
- One copper atom on average has an atomic mass of 63.55 amu.
- 1 mol of copper atoms has a mass of 63.55 g.

(It is the same number expressed in grams rather than amu)

- The molar mass of copper is 63.55 g/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 g sulfur = 1 mol sulfur = 6.022×10^{23} S atoms

12.01 g carbon = 1 mol carbon = 6.022×10^{23} C atoms

6.94 g lithium = 1 mol lithium = 6.022×10^{23} Li atoms

- The lighter the atom, the less mass in 1 mol of that atom.

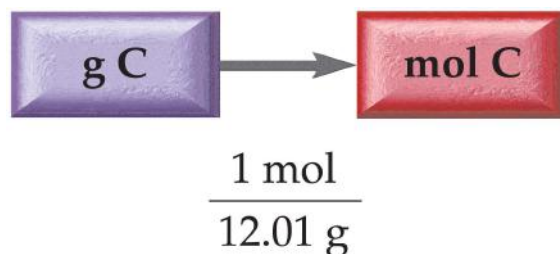
Converting Between Grams and Moles:

Calculate the number of moles of carbon in 0.58-g diamond.

GIVEN: 0.58 g C

FIND: mol C

SOLUTION MAP



RELATIONSHIPS USED

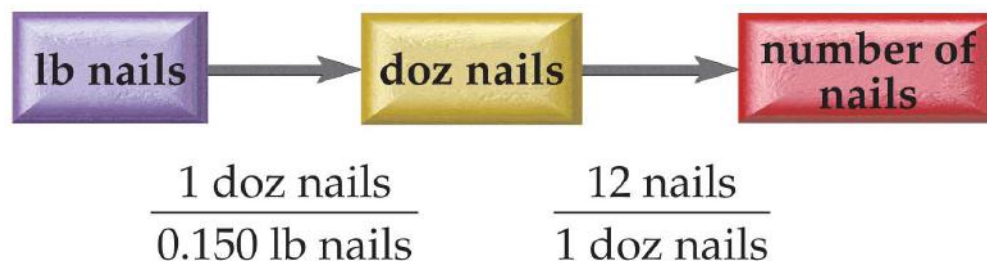
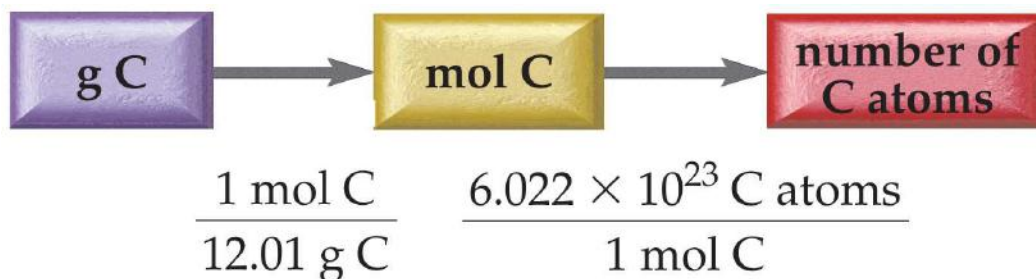
12.01 g C = 1 mol C (molar mass of carbon, from periodic table)

SOLUTION

$$0.58 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.8 \times 10^{-2} \text{ mol C}$$

Converting Grams to Moles to Atoms:

Calculate the number of atoms of carbon in 0.58-g diamond.



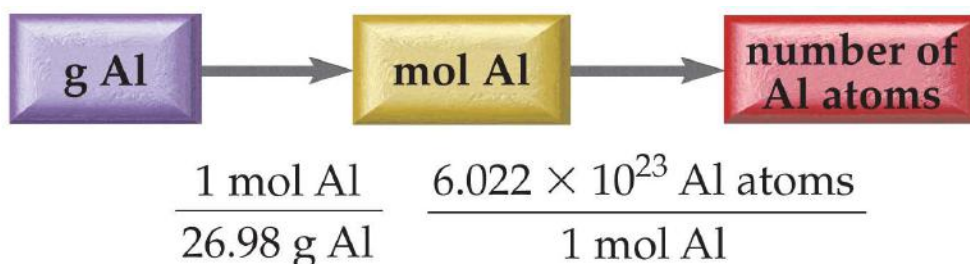
$$0.58 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C}} = 2.9 \times 10^{22} \text{ C atoms}$$

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



RELATIONSHIPS USED

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

$6.022 \times 10^{23} = 1 \text{ mol}$ (Avogadro's number)

SOLUTION

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

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, and then we calculate the number of molecules (or formula units) from moles.

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

- 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.

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

GIVEN: 1.75 mol H₂O

FIND: g H₂O

SOLUTION MAP



$$\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}}$$

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

RELATIONSHIPS USED:=

$$\begin{aligned} \text{H}_2\text{O molar mass} &= 2(\text{Atomic mass H}) + 1(\text{Atomic mass O}) \\ &= 2(1.01) + 1(16.00) \\ &= 18.02 \text{ g/mol} \end{aligned}$$

SOLUTION

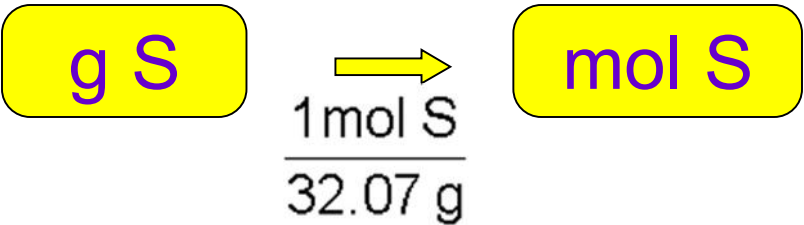
$$1.75 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}} = 31.5 \text{ g H}_2\text{O}$$

Practice:

Check your solution on next page

Calculate the moles of sulfur in 57.8 g of sulfur

Practice — Calculate the moles of sulfur in 57.8 g of sulfur

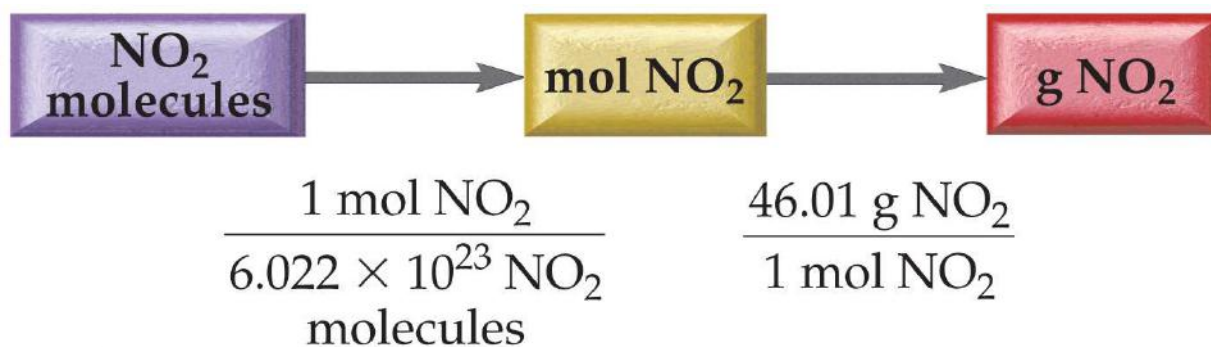
Given: Find:	57.8 g S mol S
Conceptual Plan:	
Relationships:	1 mol S = 32.07 g
Solution:	$57.8 \cancel{\text{g S}} \times \frac{1 \text{ mol}}{32.07 \cancel{\text{g}}} = 1.80 \text{ mol S}$
Check:	because the given amount is much less than 1 mol S, the number makes sense

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



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)

NO_2 molar mass = 1(Atomic mass N) + 2(Atomic mass O)
= 14.01 + 2(16.00)
= 46.01 g/mol

SOLUTION

$$4.78 \times 10^{24} \text{NO}_2 \text{ molecules} \times \frac{1 \text{ mol NO}_2}{6.022 \times 10^{23} \text{NO}_2 \text{ molecules}} \times \frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2} = 365 \text{ g NO}_2$$

Practice:

Check your solution on next page

How many copper atoms are in a penny weighing 3.10 g?

Practice: How many copper atoms are in a penny weighing 3.10 g?

Given: Find:	3.10 g Cu atoms Cu
Conceptual Plan:	<p style="text-align: center;"> $\text{g Cu} \xrightarrow{\frac{1 \text{ mol}}{63.55 \text{ g}}} \text{mol Cu} \xrightarrow{\frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}}} \text{atoms Cu}$ </p>
Relationships:	1 mol Cu = 63.55 g, 1 mol = 6.022×10^{23}
Solution:	$3.10 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}}$ $= 2.94 \times 10^{22} \text{ atoms Cu}$
Check:	because the given amount is much less than 1 mol Cu, the number makes sense

Practice:

Check your solution on next page

- Find the number of CO_2 molecules in 10.8 g of dry ice

Practice: Find the number of CO₂ molecules in 10.8 g of dry ice

Given:	10.8 g CO ₂
Find:	molecules CO ₂
Conceptual Plan:	$\text{g CO}_2 \xrightarrow[\frac{1 \text{ mol}}{44.01 \text{ g}}]{} \text{mol CO}_2 \xrightarrow[\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}}]{} \text{molec CO}_2$
Relationships:	<p>1 mol CO₂ = 44.01 g, 1 mol = 6.022 × 10²³</p>
Solution:	$10.8 \cancel{\text{g CO}_2} \times \frac{1 \cancel{\text{mol CO}_2}}{44.01 \cancel{\text{g CO}_2}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \cancel{\text{mol}}}$ $= 1.48 \times 10^{23} \text{ molecules CO}_2$
Check:	because the given amount is much less than 1 mol CO ₂ , the number makes sense

Practice:

Check your solution on next page

How many formula units are in 50.0 g of PbO_2 ? ($\text{PbO}_2 = 239.2$)

Check your solution on the next slide

Practice — How many formula units are in 50.0 g of PbO_2 ? ($\text{PbO}_2 = 239.2$)

Given:	50.0 g PbO_2
Find:	formula units PbO_2
Conceptual Plan:	$\text{g PbO}_2 \xrightarrow{\frac{1 \text{ mol}}{239.2 \text{ g}}} \text{mol PbO}_2 \xrightarrow{\frac{6.022 \times 10^{23} \text{ units}}{1 \text{ mol}}} \text{units PbO}_2$
Relationships:	1 mol $\text{PbO}_2 = 239.2 \text{ g}$, 1 mol = 6.022×10^{23}
Solution:	$50.0 \text{ g PbO}_2 \times \frac{1 \text{ mol PbO}_2}{239.2 \text{ g PbO}_2} \times \frac{6.022 \times 10^{23} \text{ units}}{1 \text{ mol}}$ $= 1.26 \times 10^{23} \text{ units PbO}_2$
Check:	because the given amount is less than 1 mol PbO_2 , the number makes sense

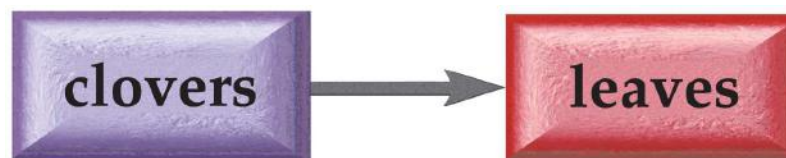
Chemical Formulas as Conversion Factors

3-Leaf Clover Analogy: How Many Leaves on 14 Clovers?

3 leaves : 1 clover



3 leaves : 1 clover

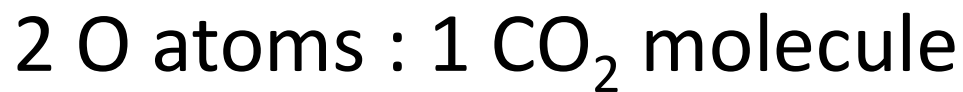


$\frac{3 \text{ leaves}}{1 \text{ clover}}$

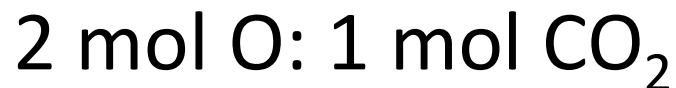
$$14 \text{ clovers} \times \frac{3 \text{ leaves}}{1 \text{ clover}} = 42 \text{ leaves}$$

Chemical Formulas as Conversion Factors

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



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



The Conversion Factor Comes Directly from the Chemical Formula

8 legs : 1 spider



4 legs : 1 chair



2 H atoms : 1 H₂O molecule



Converting Between Moles of a Compound

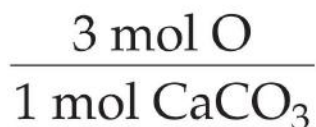
and Moles of a Constituent Element:

Find the number of moles of O in 1.7 mol CaCO_3 .

GIVEN: 1.7 mol CaCO_3

FIND: mol O

SOLUTION MAP



RELATIONSHIPS USED

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

SOLUTION

$$1.7 \text{ mol } \cancel{\text{CaCO}_3} \times \frac{3 \text{ mol O}}{1 \cancel{\text{ mol CaCO}_3}} = 5.1 \text{ mol O}$$

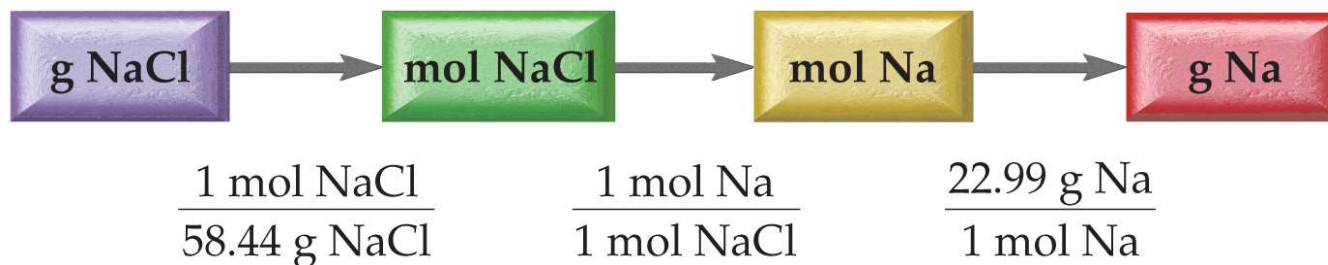
Converting Between Grams of a Compound and Grams of a Constituent Element:

Find the mass of sodium in 15 g of NaCl.

GIVEN: 15 g NaCl

FIND: g Na

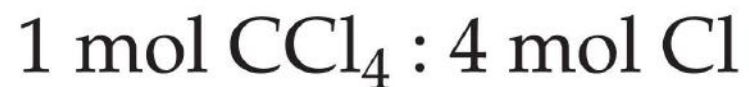
SOLUTION MAP



SOLUTION

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

Mole Relationships from a Chemical Formula



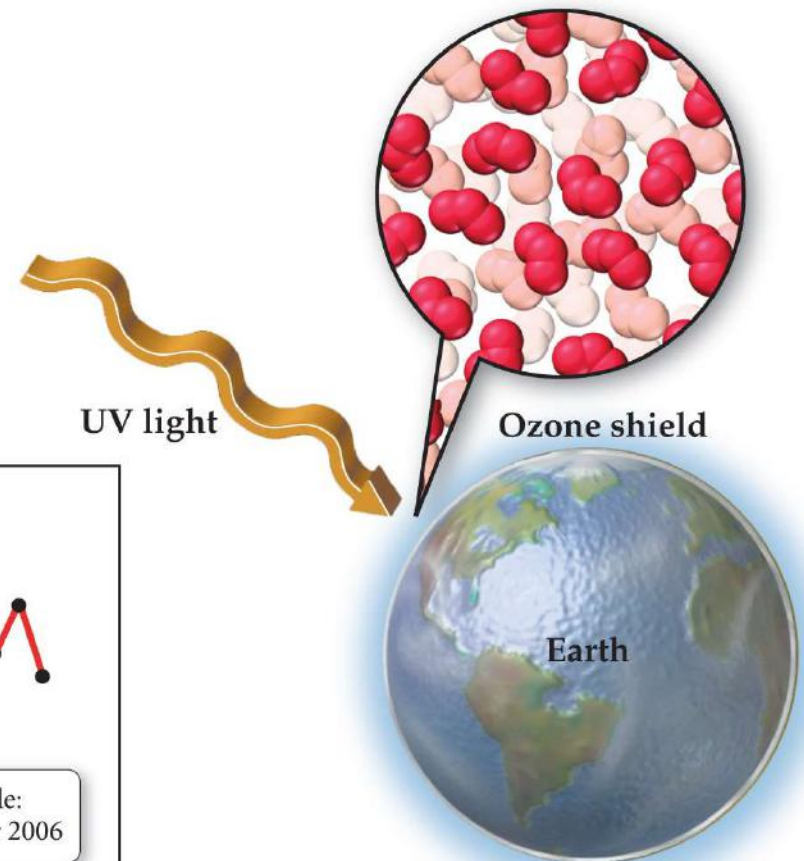
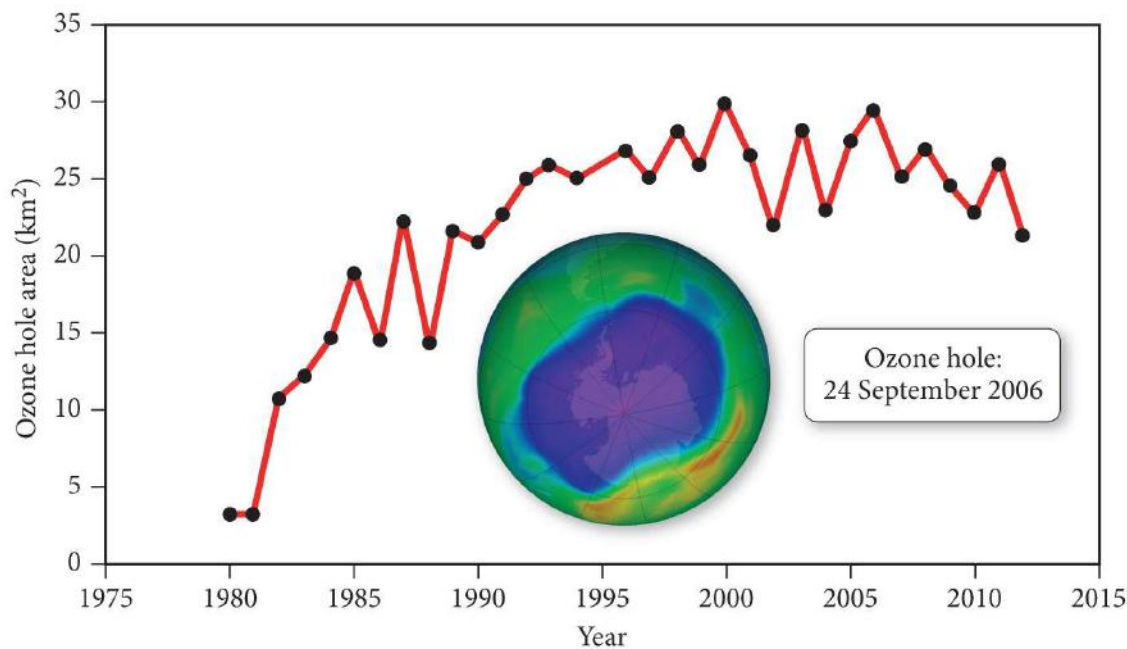
The relationships inherent in a chemical formula allow us to convert between moles of the compound and moles of a constituent element (and vice versa).

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.

Chemistry in the Environment: The Ozone Shield

- Upper atmospheric ozone is important because it acts as a shield to protect life on Earth from harmful ultraviolet light.
- Antarctic ozone hole area from 1980 to 2012. The darkest blue colors indicate the lowest ozone levels.



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.

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

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 follows:

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.

How Much Sodium in Sodium Chloride?

GIVEN:

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

FIND: g NaCl

SOLUTION MAP



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

RELATIONSHIPS USED

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

SOLUTION

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

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.

Practice:

Check your solution on next page

Benzaldehyde is 79.2% carbon. What mass of benzaldehyde contains 19.8 g of C?

Practice — Benzaldehyde is 79.2% carbon. What mass of benzaldehyde contains 19.8 g of C?

Given:	19.8 g C, 79.2% C
Find:	g benzaldehyde
Conceptual Plan:	
Relationships:	100 g benzaldehyde : 79.2 g C
Solution:	$19.8 \text{ g C} \times \frac{100 \text{ g benzaldehyde}}{79.2 \text{ g C}} = 25.0 \text{ g benzaldehyde}$
Check:	because the mass of benzaldehyde is more than the mass of C, the number makes sense

Practice:

Check your solution on next page

How many grams of sodium are in 6.2 g of NaCl? (Na = 22.99; Cl = 35.45)



Practice — Find the mass of sodium in 6.2 g of NaCl

Given: Find:	6.2 g NaCl g Na
Conceptual Plan:	$ \begin{array}{ccccccc} \text{g NaCl} & \longrightarrow & \text{mol NaCl} & \longrightarrow & \text{mol Na} & \longrightarrow & \text{g Na} \\ \frac{1 \text{ mol}}{58.44 \text{ g}} & & \frac{1 \text{ mol Na}}{1 \text{ mol NaCl}} & & \frac{22.99 \text{ g}}{1 \text{ mol}} & & \\ \text{Na} = 1 \times \frac{22.99 \text{ amu}}{35.45 \text{ amu}} & & & & & & \\ \text{NaCl} = & & & & & & 58.44 \text{ amu} \end{array} $
Relationships:	1 mol NaCl = 58.44 g, 1 mol Na = 22.99 g, 1 mol Na : 1 mol NaCl $\text{Na} = 1 \times \frac{22.99 \text{ amu}}{35.45 \text{ amu}}$ $\text{NaCl} = 58.44 \text{ amu}$
Solution:	$ 6.2 \cancel{\text{g NaCl}} \times \frac{1 \cancel{\text{mol NaCl}}}{58.44 \cancel{\text{g}}} \times \frac{1 \cancel{\text{mol Na}}}{1 \cancel{\text{mol NaCl}}} \times \frac{22.99 \text{ g Na}}{1 \cancel{\text{mol Na}}} = 2.4 \text{ g Na} $
Check:	because the amount of Na is less than the amount of NaCl, the answer makes sense

Mass Percent Composition from a Chemical Formula

Based on the chemical formula, the mass percent of element Cl in compound CCl_2F_2 is as follows:

CCl_2F_2

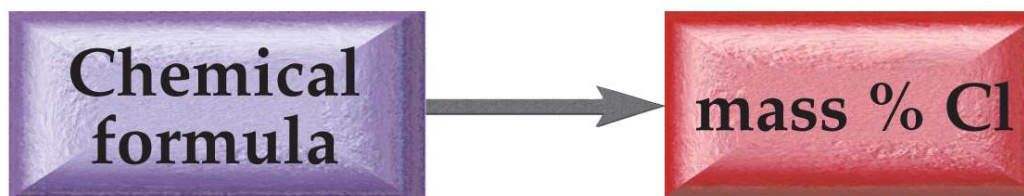

$$\text{Mass percent Cl} = \frac{2 \times \text{Molar mass Cl}}{\text{Molar mass CCl}_2\text{F}_2} \times 100\%$$


**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



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

Mass Percent Composition from a Chemical Formula: Calculate the mass percent of Cl in $C_2Cl_4F_2$, freon-114.

RELATIONSHIPS USED

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

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

$$\text{Molar mass } C_2Cl_4F_2 = 2(12.01) + 4(35.45) + 2(19.00)$$

$$= 24.02 + 141.8 + 38.00$$

$$= \frac{203.8 \text{ g}}{\text{mol}}$$

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

$$= \frac{141.8 \text{ g}}{203.8 \text{ g}} \times 100\%$$

$$= 69.58\%$$

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 as follows:

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

Practice:

Check your solution on next page

Determine the mass percent composition of the following

CaCl_2 (Ca = 40.08, Cl = 35.45)

Practice — Determine the percent composition of the following

- CaCl_2 $\text{Mass \% Ca} = \frac{\text{molar mass Ca}}{\text{molar mass CaCl}_2} \times 100\%$
 $\text{Mass \% Cl} = \frac{2 \times \text{molar mass Cl}}{\text{molar mass CaCl}_2} \times 100\%$

$$2 \times \text{molar mass Cl} = 2(35.45 \text{ g/mol}) = 70.90 \text{ g/mol}$$

$$\text{molar mass CaCl}_2 = 1(40.08) + 2(35.45) = 110.98 \text{ g/mol}$$

$$\text{Mass \% Ca} = \frac{40.08 \text{ g/mol}}{110.98 \text{ g/mol}} \times 100\% = 36.11\%$$

$$\text{Mass \% Cl} = \frac{70.90 \text{ g/mol}}{110.98 \text{ g/mol}} \times 100\% = 63.88\%$$

Practice:

Check your solution on next page

- Find the % of Nitrogen in Ammonium Phosphate?

Practice

- Find the % of Nitrogen in Ammonium Phosphate?

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

Ammonium Phosphate is $(\text{NH}_4)_3 \text{PO}_4$

$$\%N = (14\text{g} \times 3 / 146\text{g}) \times 100\% = 28.8\%$$

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.

Practice:

Check your solution on next page

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?

Practice Chemistry and Health: Fluoridation of Drinking Water

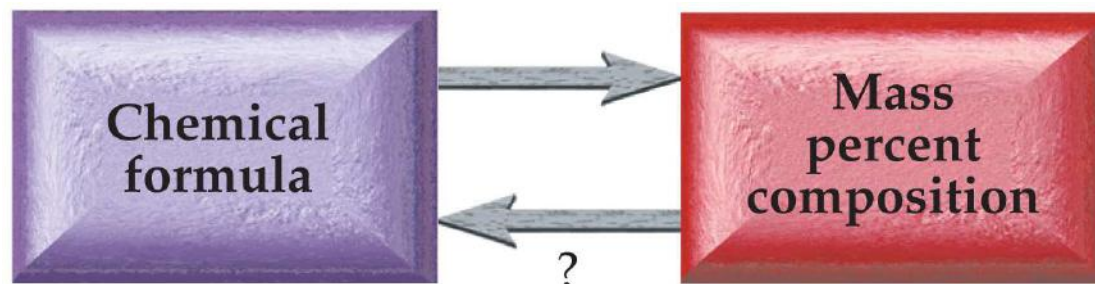
- What is the mass percent composition of F^- in NaF?

$$(19\text{g}/42\text{g}) \times 100\% = 45.2\%$$

- 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?

$$1500\text{L} \times (1\text{mg}F^-/\text{L}) (10^{-3}\text{g}F^-/1\text{mg}F^-) (100\text{gNaF}/45.2\text{g}F^-) = 3.3\text{g}$$

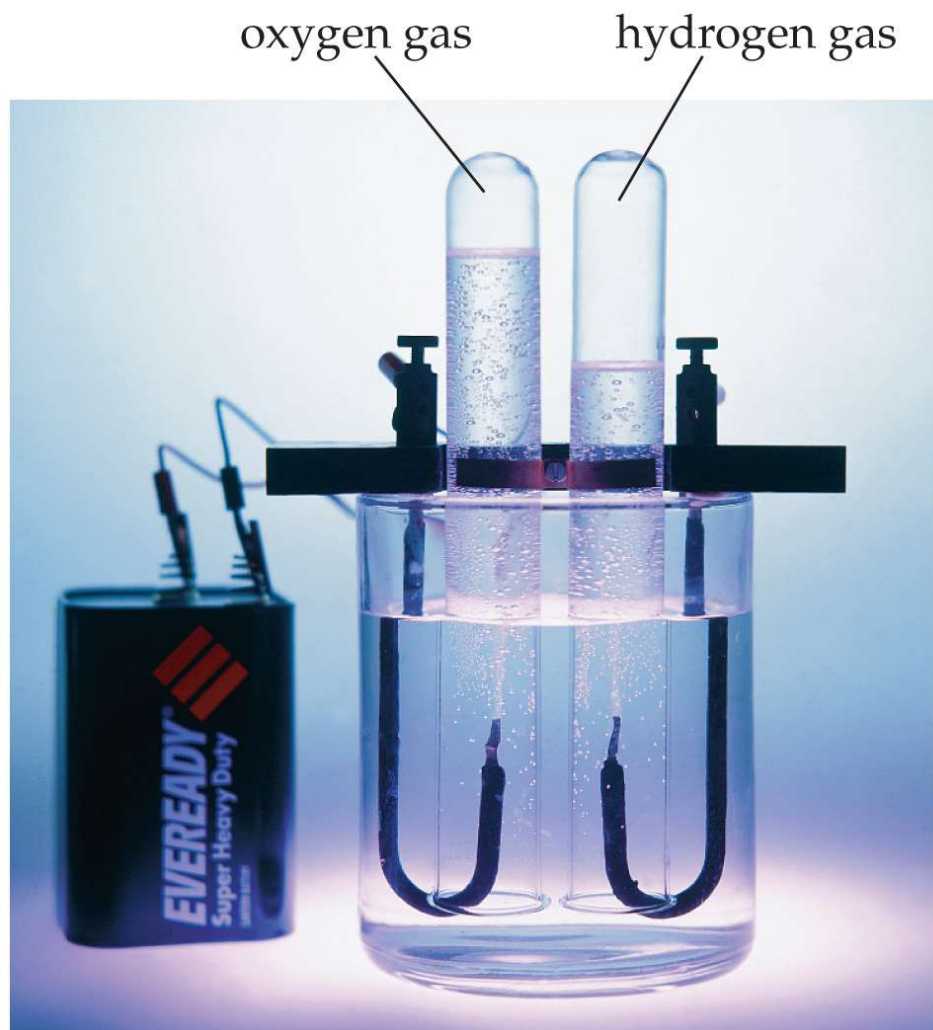
Empirical Formulas from Mass Percent Composition



- An empirical formula gives only 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.
- Molecular formula = Empirical $\times n$, where $n = 1, 2, 3 \dots$
- $n = 2$ for hydrogen peroxide

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

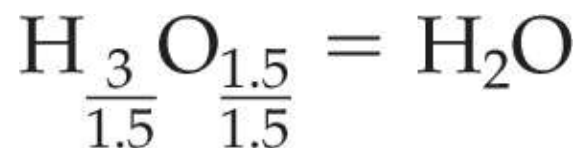
$$\text{Moles H} = 3.0 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 3.0 \text{ mol H}$$

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

- 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:
 $\text{H}_3\text{O}_{1.5}$
- To get whole-number subscripts in our formula, divide all the subscripts by the smallest one, in this case 1.5.



- 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.

Conversion of Fractional Subscripts to Whole Numbers

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	$\frac{1}{10}$	10
_.20	$\frac{1}{5}$	5
_.25	$\frac{1}{4}$	4
_.33	$\frac{1}{3}$	3
_.50	$\frac{1}{2}$	2
_.66	$\frac{2}{3}$	3
_.75	$\frac{3}{4}$	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.

$$\text{Mass Ti} = 3.24 \text{ g Ti}$$

$$\begin{aligned}\text{Mass O} &= \text{Mass oxide} - \text{Mass titanium} \\ &= 5.40 \text{ g Ti and O} - 3.24 \text{ g Ti} \\ &= 2.16 \text{ g O}\end{aligned}$$

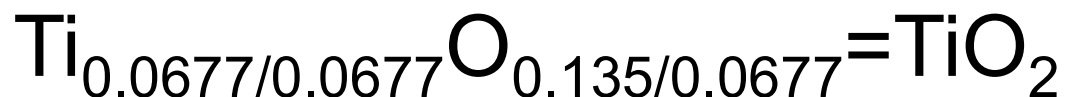
- Now you can convert the mass of each element to moles.

Calculating an Empirical Formula from Reaction Data

- $3.24 \text{ g Ti} \times (1 \text{ mol Ti} / 47.88 \text{ g}) = 0.0677 \text{ mol Ti}$
- $2.16 \text{ g O} \times (1 \text{ mol O} / 16.00 \text{ g}) = 0.135 \text{ mol O}$



We need whole numbers



Calculating Molecular Formulas for Compounds from Empirical Formulas and Molar Masses

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

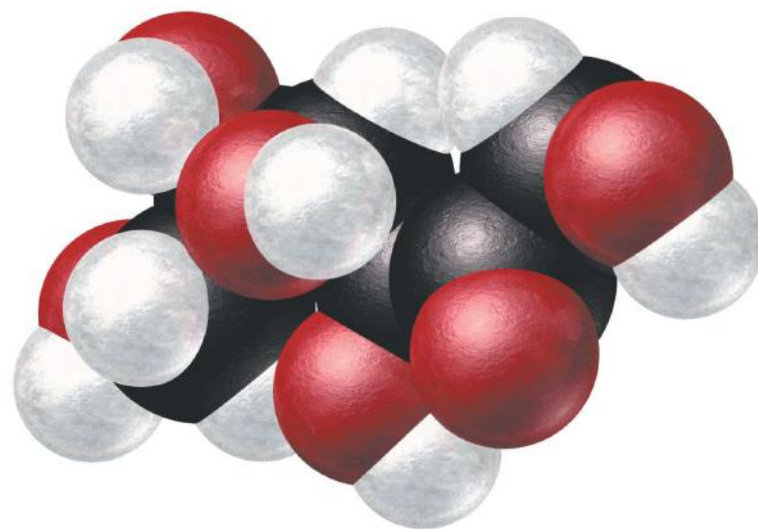


- We can find n in the expression
Molar mass = Empirical formula molar mass $\times n$
- Solving for n ,

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

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 .



Calculating Molecular Formulas for Compounds: Fructose

- For fructose, the empirical formula molar mass is as follows:

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.

$$\text{Molecular formula} = \text{CH}_2\text{O} \times 6 = \text{C}_6\text{H}_{12}\text{O}_6$$

Calculating Molecular Formulas for Compounds -- if we are given the empirical formula already

- 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.

Calculating Molecular Formulas for Compounds

-- if we are given the % mass composition and molar mass it might be easier to do the following:

- Multiply molar mass by the mass fraction of each element
- Convert mass of each element to moles (it should automatically be a whole number)
- Number of moles of each element is the subscript in the molecular formula
- If you need to find the empirical formula, simplify the subscripts in the molecular formula to the smallest set of whole numbers

Example:
**Find the empirical
formula of aspirin
with the given
mass percent
composition**

Write down the given quantity and its units

Given: C = 60.00%

H = 4.48%

O = 35.53%

Therefore, in 100 g of aspirin there are 60.00 g C,
4.48 g H, and 35.53 g O

Example:
**Find the empirical
formula of aspirin
with the given
mass percent
composition**

Information

Given: 60.00 g C, 4.48 g H, 35.53 g O

Write down the quantity to find and/or its units

Find: empirical formula, $C_xH_yO_z$

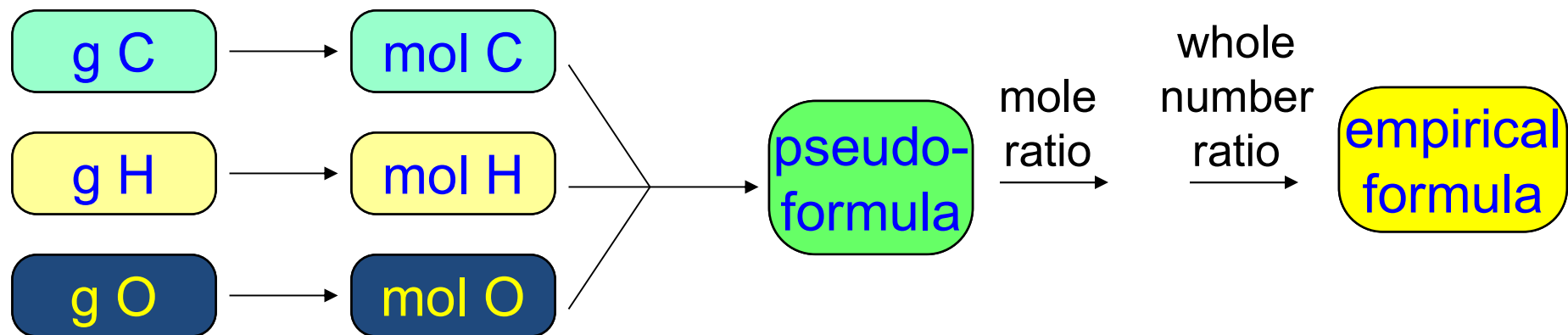
Example:
Find the empirical formula of aspirin with the given mass percent composition

Information

Given: 60.00 g C, 4.48 g H, 35.53 g O

Find: empirical formula, $C_xH_yO_z$

- Write a conceptual plan



Example:
**Find the empirical
formula of aspirin
with the given
mass percent**

Information

Given: 60.00 g C, 4.48 g H, 35.53 g O

Find: empirical formula, $C_xH_yO_z$

CP: g C,H,O \rightarrow mol C,H,O \rightarrow

pseudo form. \rightarrow mol ratio \rightarrow emp. form.

composition

Collect needed relationships

1 mole C = 12.01 g C

1 mole H = 1.008 g H

1 mole O = 16.00 g O

**Example:
Find the empirical
formula of aspirin
with the given
mass percent
composition**

Information

Given: 60.00 g C, 4.48 g H, 35.53 g O

Find: empirical Formula, $C_xH_yO_z$

CP: g C,H,O \rightarrow mol C,H,O \rightarrow

pseudo form. \rightarrow mol ratio \rightarrow emp. form.

Rel: 1 mol C = 12.01 g; 1 mol H = 1.008 g;

1 mol O = 16.00 g

Apply the conceptual plan

- calculate the moles of each element

$$60.00 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.996 \text{ mol C}$$

$$4.48 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.44 \text{ mol H}$$

$$35.53 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.220 \text{ mol O}$$

Example:
Find the empirical formula of aspirin with the given mass percent composition

Information

Given: 60.00 g C, 4.48 g H, 35.53 g O

Find: empirical formula, $C_xH_yO_z$

CP: g C,H,O \rightarrow mol C,H,O \rightarrow

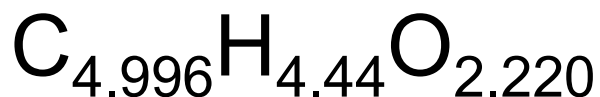
pseudo form. \rightarrow mol ratio \rightarrow emp. form.

Rel: 1 mol C = 12.01 g; 1 mol H = 1.008 g;

1 mol O = 16.00 g

Apply the conceptual plan

– write a pseudoformula



**Example:
Find the empirical
formula of aspirin
with the given
mass percent
composition**

Information

Given: 60.00 g C, 4.48 g H, 35.53 g O

Find: empirical formula, $C_xH_yO_z$

CP: g C,H,O \rightarrow mol C,H,O \rightarrow

pseudo form. \rightarrow mol ratio \rightarrow emp. form.

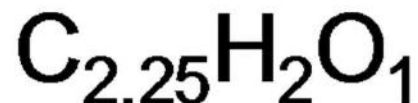
Rel: 1 mol C = 12.01 g; 1 mol H = 1.008 g;

1 mol O = 16.00 g

Apply the concept plan

- find the mole ratio by dividing by the smallest number of moles

$$\begin{array}{ccc} \text{C} & \text{H} & \text{O} \\ 4.996 & 4.44 & 2.220 \\ \hline 2.220 & 2.220 & 2.220 \end{array}$$



Example:
Find the empirical formula of aspirin with the given mass percent composition

Information

Given: 60.00 g C, 4.48 g H, 35.53 g O

Find: empirical formula, $C_xH_yO_z$

CP: g C,H,O \rightarrow mol C,H,O \rightarrow

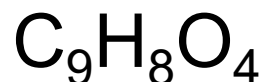
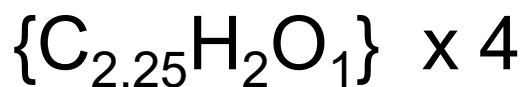
pseudo form. \rightarrow mol ratio \rightarrow emp. form.

Rel: 1 mol C = 12.01 g; 1 mol H = 1.008 g;

1 mol O = 16.00 g

Apply the conceptual plan

- multiply subscripts by factor to give whole number



Practice:

Check your solution on next page

Determine the empirical formula of stannous fluoride, which contains 75.7% Sn (118.70g/mol) and the rest fluorine (19.00g/mol)

Practice — Determine the empirical formula of stannous fluoride, which contains 75.7% Sn (118.70) and the rest fluorine (19.00)

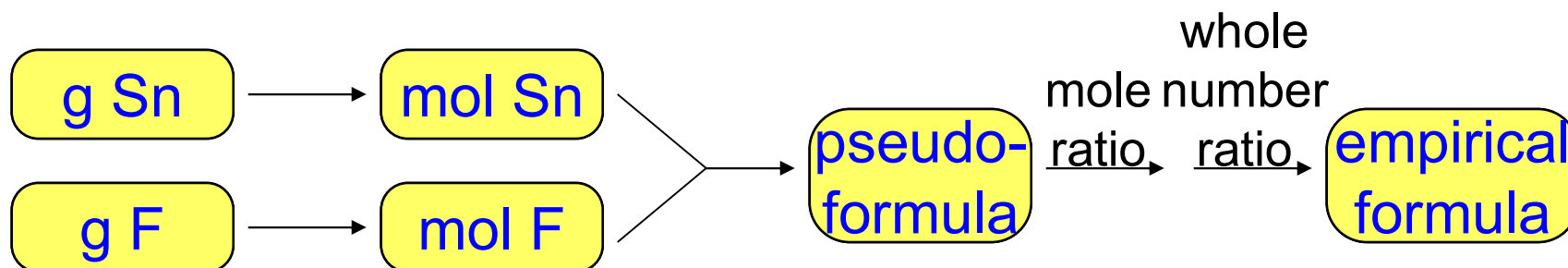
Given: 75.7% Sn, $(100 - 75.3) = 24.3\%$ F \therefore

in 100 g stannous fluoride there are 75.7 g Sn and 24.3 g F

Find: Sn_xF_y

Rel: 1 mol Sn = 118.70 g; 1 mol F = 19.00 g

Conceptual plan:



Practice — Determine the empirical formula of stannous fluoride, which contains 75.7% Sn (118.70g/mol) and the rest fluorine (19.00g/mol)

Apply the conceptual plan

$$75.7 \text{ g Sn} \times \frac{1 \text{ mol Sn}}{118.70 \text{ g}} = 0.638 \text{ mol Sn}$$

$$24.3 \text{ g F} \times \frac{1 \text{ mol F}}{19.00 \text{ g}} = 1.28 \text{ mol F}$$



Practice:

Check your solution on next page

Determine the empirical formula of magnetite, which contains 72.4% Fe (55.85g/mol) and the rest oxygen (16.00g/mol)

Practice — Determine the empirical formula of magnetite, which contains 72.4% Fe (55.85g/mol) and the rest oxygen (16.00g/mol)

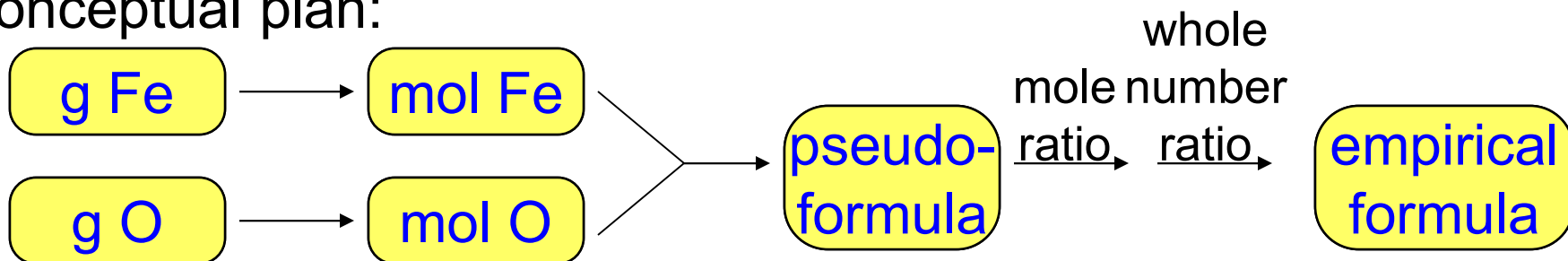
Given: 72.4% Fe, $(100 - 72.4) = 27.6\%$ O \therefore

in 100 g magnetite there are 72.4 g Fe and 27.6 g O

Find: Fe_xO_y

Rel: 1 mol Fe = 55.85 g; 1 mol O = 16.00 g

Conceptual plan:

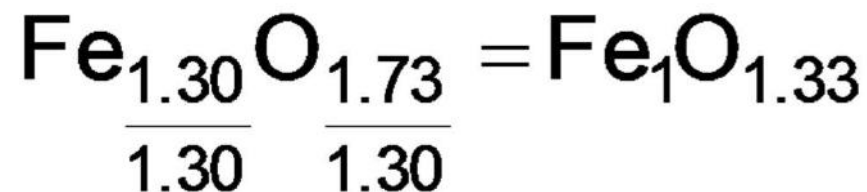
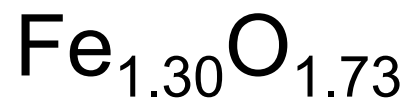


Practice — Determine the empirical formula of magnetite, which contains 72.4% Fe (55.85g/mol) and the rest oxygen (16.00g/mol)

Apply the conceptual plan

$$72.4 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g}} = 1.30 \text{ mol Fe}$$

$$26.7 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g}} = 1.73 \text{ mol O}$$



Chapter 6 in Review

The Mole Concept:

- The mole is a specific number (6.022×10^{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.

Practice

- Laboratory analysis of aspirin determined the following mass percent composition. Find the empirical formula.

$$\text{C} = 60.00\%$$

$$\text{H} = 4.48\%$$

$$\text{O} = 35.53\%$$

Find the empirical formula of aspirin with the given mass percent composition

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 Learning Objectives

1. LO: Convert between moles and number of atoms.
2. LO: Convert between grams and moles.
3. LO: Convert between grams and number of atoms or molecules.
4. LO: Convert between moles of a compound and moles of a constituent element.
5. LO: Convert between grams of a compound and grams of a constituent element.
6. LO: Use mass percent composition as a conversion factor.
7. LO: Determine mass percent composition from a chemical formula.
8. LO: Determine an empirical formula from experimental data.
9. LO: Calculate a molecular formula from an empirical formula and molar mass (or an empirical formula from the molecular formula)