Chapter 6: Chemical Composition

Read Chapter 6

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How Much?

Chemical composition along with atomic and formula masses are the keys to finding the answer to “How much…?” Knowing the chemical composition, formula mass, mass percentage of each element, number of grams, particles, moles, ratios etc. will enable us to predict outcomes.

Avogadro’s Number:

Just as a dozen is the term that represents 12 pieces (a package of 1 dozen is equivalent to 12 eggs, 12 donuts, 12 long stem roses), one MOLE represents a much larger amount $6.022 \times 10^{23}$ pieces. This number was chosen because it will take that many atoms of an element to have the mass in grams equivalent to the weighted average mass number on the periodic table. This number is huge. If we had $6.022 \times 10^{23}$ marbles, they would cover the entire earth 50 feet high, yet this number of water molecules will have a mass of 18.0 grams or a volume of 18.0 milliliters. Atoms are very tiny.

1 mol = 1 mole = $6.022 \times 10^{23}$ particles… these particles are generally atoms, ions, molecules, or formula units.

Twenty-two copper pennies (pre 1982) contain about 1 mole of copper atoms
Stoichiometry:  Chemical arithmetic.

Stoichiometry is a fancy name to represent the math and conversions used in converting mass to moles, molecules to moles, moles of one compound to moles of another in a balanced chemical reaction, and many other math conversions related to chemicals amounts.

Stoichiometry Equations and Examples

Molar Mass:  Recall that the mass of 1 mole of atoms of an element is its molar mass or atomic weight or atomic mass; 1 mole of molecules is its molar mass or molecular weight or molecular mass; one mole of a compound (formula unit) is its molar mass or formula weight or formula mass.  Mass values are found on the Periodic Table.

molar mass (MM):  atomic mass, atomic weight, formula weight, molecular weight (MW), formula mass, molecular mass, element mass.

1. Finding the average atomic mass of an element:
look on the periodic table.
Example 1:
atomic mass of Na = 22.99  or atomic mass of Cl = 35.45
The units we use are usually grams/mole (amu/atom is also true).
1 amu = 1.66 x 10^{-24} g (amu is atomic mass unit)

2. Solve for molar mass:
Molecular weight (MW) = the summation of the number of atoms of each element times their atomic mass.  Always keep a minimum of one place past the decimal.
Example 2:
What is the molar mass of Al$_2$S$_3$?
2 Al = 2(26.98) = 53.96 g/mole
3 S = 3(32.06) = 96.18 g/mole
MW of Al$_2$S$_3$ = 150.14 g/mole
Try this 2:
a) What is the molar mass of Cu?

b) What is the molar mass of table sugar, C$_{12}$H$_{22}$O$_{11}$?
Converting between moles and number of atoms/molecules/formulas:

3. To find the number of atoms/molecules/formula units from known moles:
   molecules = moles x 6.022 x 10^{23} molecules (Avogadro's number)

   Example 3:
   How many formula units are in 3.20 moles of Al_2S_3?
   (3.20 moles of Al_2S_3) x (6.022 x 10^{23} formula units/mole)
   = 1.93 x 10^{24} formula units of Al_2S_3.

   Try this 3:
   How many molecules are in 12.8 moles of table sugar, C_{12}H_{22}O_{11}?

Converting between atoms and molecule/formula units

4. To find the number of atoms of a particular element in a given number of molecules:
   atoms = molecules x the number of atoms in each molecule

   Example 4:
   How many Al atoms are in 1.93 x 10^{24} molecules of Al_2S_3?
   (1.93 x 10^{24} molecules Al_2S_3) x (2 Al atoms/1 molecule Al_2S_3) = 3.86 x 10^{24} atoms Al

   Try this 4:
   a) How many H atoms are in 12.8 moles of table sugar, C_{12}H_{22}O_{11}?

   b) How many total atoms are in 12.8 moles of table sugar, C_{12}H_{22}O_{11}?
Converting between grams and moles

5. To find the mass of a compound and number of moles are known:
   \[ \text{mass} = \text{MM} \times \text{moles} \]

   **Example 5:**
   What is the mass of 3.20 moles of \( \text{Al}_2\text{S}_3 \)? (from example 2 we know the MM is 150.14 g/mole)
   
   \[150.14 \text{ g/mole} \times 3.20 \text{ moles} = 480. \text{ grams} \]

   **Try this 5:**
   What is the mass of 12.8 moles of table sugar, \( \text{C}_{12}\text{H}_{22}\text{O}_{11} \)?

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6. To find the number of moles when mass is known:
   \[ \text{moles} = \frac{\text{mass}}{\text{MM}} \]

   **Example 6.**
   How many moles of \( \text{Al}_2\text{S}_3 \) are contained in 564 grams?

   (MW still is 150.14 g/mol)
   
   \[ \frac{564\text{g}}{150.14\text{g/mol}} = 3.76 \text{ moles of Al}_2\text{S}_3. \]

   **Try this 6:**
   How many moles of table sugar, \( \text{C}_{12}\text{H}_{22}\text{O}_{11} \), are contained in 12.0 grams?
Now try to combine what you learned.

a) Calculate the number of grams of just the sodium in 25.0 grams of NaCl,

\[
\begin{align*}
g \text{ NaCl} & \rightarrow \text{ mol NaCl} \\
& \rightarrow \text{ mol Na} \\
& \rightarrow g \text{ Na}
\end{align*}
\]

\[
\begin{align*}
1 \text{ mol NaCl} & \rightarrow 58.44 g \text{ NaCl} \\
1 \text{ mol Na} & \rightarrow 1 \text{ mol NaCl} \\
22.99 g \text{ Na} & \rightarrow 1 \text{ mol Na}
\end{align*}
\]

b) Calculate the number of grams of carbon in 12.0 grams of table sugar, C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}

Mass Percent Composition:

7. To find a mass percent of an element in a given compound:
   mass % = \left( \frac{\text{mass of element in 1 mole of compound}}{\text{MM}} \right) \times 100\%

Example 7:
What is the mass percent of S in Al\textsubscript{2}S\textsubscript{3}?

The molar mass is 150.14 g/mole; the mass of S is 3(32.06) = 96.18 g/mole

\[\left( \frac{96.18 \text{ g/mol S}}{150.14 \text{ g/mol Al}_2S_3} \right) \times 100\% = 64.06\%
\]

Try this 7:
What is the mass percent of each component in table sugar, C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}?
Empirical Formula

8. Empirical formulas are the simplest whole-number ratio of atoms in a compound. To get the empirical formula divide the formula subscripts by the lowest common multiple.

Example 8:
C₆H₆ divide the subscript numbers by 6 = CH
C₈H₆O₄ divide the subscript numbers by 2 = C₄H₃O₂

Try this 8:
Solve for the empirical formula of each.
a) Butane, C₄H₁₀

b) Hydrogen peroxide, H₂O₂
c) Benzene, C₆H₁₂
d) Aluminum sulfide, Al₂S₃

9. To calculate the molecular formula when the empirical formula and the molar mass are known:
Molar mass / empirical formula mass = n  MM/EM = n
multiply subscripts in the empirical formula by n

Example 9:
A compound has an empirical formula of NO₂ and is found to have a molar mass of 92g. What is the molecular formula? The empirical formula mass = 46.0g
92g total/46.0g NO₂ = 2
molecular formula = N₂O₄

Try this 9:
A compound has an empirical formula of C₂H₄O and is found to have a molar mass of 176 g. What is the molecular formula?
10. To calculate empirical formulas when you know the mass of each element in the compound,
   
   1. If given % by mass; change the % sign to grams
      ( 64.1% = 64.1g)
   
   2. Convert the mass of each element to moles (example 6)
   
   3. Divide the number of moles of each element by the smallest number of moles
      to convert the smallest number to 1.
   
   4. If all mole numbers are now integers, use these numbers as the subscripts in the
      empirical formula.
   
   5. If the numbers are not within 0.1 of a whole number, then multiply the
      numbers by the smallest integer necessary (such as 2,3,4,5) to convert them
      all to whole numbers. Use whole numbers as subscripts

   \[ \text{Decimal} \quad \text{Multiply by} \]
   \[ .50 \quad 2 \]
   \[ .33 \text{ or } .67 \quad 3 \]
   \[ .25 \text{ or } .75 \quad 4 \]
   \[ .20, .40, .60, .80 \quad 5 \]

   Example 10:
   
   A compound was analyzed to have 11.98g Al and 21.35g of sulfur. What is the
   empirical formula?
   
   1. we have grams
   
   2. convert to moles using example 6 method (moles = mass/MW)
      atomic mass Al = 26.98g/mole; atomic mass of S = 32.06g/mole
      \[
      (11.98\text{g Al}) / (26.98\text{g/mol Al}) = 0.4440 \text{ mol Al} \\
      (21.35\text{g S}) / (32.06\text{g/mol S}) = 0.6659 \text{ mol S}
      \]
   
   3. divide each by 0.4440 mol (smallest number of moles)
      0.4440mol Al / 0.4440 mol = 1 mol Al
      0.6659 mol S / 0.4440 mol = 1.50 mol S
   
   4. numbers are not all integers
   
   5. multiply each mole number by 2 to convert them to whole numbers
      1 mole Al x 2 = 2 mol Al 1.5 mol S x 2 = 3 mol S
      Use mole numbers as subscripts in the empirical number.
      Empirical formula = Al\textsubscript{2}S\textsubscript{3}
Try this 10:

a) Oil of jasmine (benzyl acetate) has 71.98% carbon, 6.71% hydrogen and 21.31% oxygen. Solve for its empirical formula.

b) Tylenol (acetaminophen) has 63.56% carbon, 6.00% hydrogen, 9.27% nitrogen and 21.17% oxygen. Solve for its empirical formula.

Coefficients

11. Coefficients (the number before the formula or element) in balanced equations:

Coefficients represent the ratio of the number of molecules or moles involved, it does NOT represent grams.

Example 11:

$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

this means it takes 2 hydrogen molecules or moles and 1 oxygen molecule or mole to produce 2 water molecules or moles.

Try this 11:

How many hydrogen molecules are needed to react with two nitrogen molecules while producing ammonia?

$\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
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.

Try this:
Suppose an old refrigerator contains 1.92 kg of freon-12 ($CCl_2F_2$).

a) How many moles of freon-12 are in that refrigerator?

b) How many molecules of freon-12 are in that refrigerator?

c) How many atoms of Cl are in it?
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.
- Fluoride is often added to water as sodium fluoride (NaF).

Try this:
   a) What is the mass percent composition of \( F^- \) in NaF?

   b) 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 Problems:

1. a) Calculate the **molar mass** of CaCl₂
   b) Calculate the number of **moles** in 4.50 g of calcium chloride
   c) Calculate the number of **formula units** in 4.50 g of CaCl₂
   d) Calculate the number of **chlorine atoms** in 4.50 g of CaCl₂
   e) Calculate the **mass** in kilograms of 1.05 x 10²⁶ formulas of CaCl₂

2. A compound consists of 2.00 g carbon, 0.333 g hydrogen and 2.66 g oxygen.
   a) Calculate the number of **moles** carbon, hydrogen and oxygen in the sample.
   b) Calculate the **empirical formula** of this compound.
   c) Given that the molecular weight is 120 g/mol, solve for the **molecular formula**?
   d) Calculate the **percent composition** of carbon, hydrogen and oxygen in this compound.

3. A sample of dimethlyhydrazine, a colorless liquid used as rocket fuel, is found to contain 39.9% carbon, 13.4% hydrogen, and 46.6% nitrogen.
   a) Calculate the **empirical formula** of the sample.
   b) If the molecular weight is 60.0 g/mol, what is its **molecular formula**?

4. a) Calculate the **molar mass** of Co₂(Cr₂O₇)₃
   b) Determine the **mass percent composition** of only Co in Co₂(Cr₂O₇)₃
   c) Calculate the number of **moles** in 21.2 g of Co₂(Cr₂O₇)₃
   d) Calculate the number of **formula units** in 21.2 g of Co₂(Cr₂O₇)₃
   e) Calculate the number of **chromium atoms** in 21.2 g of Co₂(Cr₂O₇)₃
   f) Calculate the mass of 7.25 x 10²² formula units of Co₂(Cr₂O₇)₃
   g) Name Co₂(Cr₂O₇)₃