Lesson 01: Atomic Masses and Avogadro’s Hypothesis

01 Counting Atoms and Molecules

The chemical changes we observe always involve a certain number of atoms that rearrange themselves into new configurations.

These numbers are far too large in magnitude for us to visualize let alone count, but they are still numbers, and we need to have a way to deal with them.

Because of this we need a computational bridge for these kinds of numbers.

The mole concept provides this bridge, and is central to all of quantitative chemistry.
02 Avogadro’s Number and the Mole

Avogadro's Constant or Number, is $6.022\times10^{23}$. It is one of the fundamental constants of chemistry.

It permits calculation of the amount of pure substance in chemical reactions on the basis of stoichiometric relationships or relationships between reactants used and products formed.

Avogadro’s Constant is a number, just as a dozen is number, and thus is dimensionless. You can think of Avogadro's number as the chemist's dozen. This chemist’s dozen, from this point onwards, is referred to as the mole.

$$
12 \text{ individuals} = 1 \text{ dozen individuals} \\
6.022\times10^{23} \text{ individuals} = 1 \text{ mole individuals}
$$

### Chemical Reactions and the Mole

<table>
<thead>
<tr>
<th>$2 \text{ mol } H_2$</th>
<th>$+ \quad 1 \text{ mol } O_2$</th>
<th>$= \quad 2 \text{ mol } H_2O$</th>
</tr>
</thead>
</table>


Example 01

If you have a mole or donuts, nails, dogs, atoms, molecules or electrons, how many individuals do you have?

- $6.022 \times 10^{23}$ donuts
- $6.022 \times 10^{23}$ nails
- $6.022 \times 10^{23}$ dogs
- $6.022 \times 10^{23}$ atoms
- $6.022 \times 10^{23}$ molecules
- $6.022 \times 10^{23}$ electrons
03 Atomic Mass Units (AMU) and Molar Mass

The atomic mass unit is the standard unit that is used for indicating mass on an individual atomic or molecular scale.

The relationship between atomic mass units and mass in grams is...

\[
1 \text{ atomic mass unit (u)} = \frac{\text{molar mass constant}}{\text{avogadro's constant}} = \frac{1 \text{ g particles / 1 mole particles}}{6.022 \times 10^{23} \text{ particles / 1 mole particles}} \]

\[
= \frac{1 \text{ g}}{6.022 \times 10^{23} \text{ particles}}
\]

Atomic mass units are convenient because…

• an atom or molecule that contains: \( n \) protons and neutrons
• has a mass approximately equal to: \( n \times u \)

\[
1 \text{ g of particles} = 6.022 \times 10^{23} \text{ particles}
\]

or

\[
1 \text{ g of particles} = 1 \text{ mole of particles}
\]

This is molar mass.
<table>
<thead>
<tr>
<th>Element</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Atomic Mass (u)</th>
<th>Molar Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
<td>0</td>
<td>1.00794</td>
<td>1.0079</td>
</tr>
<tr>
<td>He</td>
<td>2</td>
<td>2</td>
<td>4.002602</td>
<td>4.002602</td>
</tr>
<tr>
<td>Li</td>
<td>3</td>
<td>4</td>
<td>6.941</td>
<td>6.941</td>
</tr>
<tr>
<td>Be</td>
<td>4</td>
<td>5</td>
<td>9.012182</td>
<td>9.012182</td>
</tr>
<tr>
<td>B</td>
<td>5</td>
<td>6</td>
<td>10.811</td>
<td>10.811</td>
</tr>
<tr>
<td>C</td>
<td>6</td>
<td>6</td>
<td>12.0107</td>
<td>12.0107</td>
</tr>
<tr>
<td>N</td>
<td>7</td>
<td>7</td>
<td>14.0067</td>
<td>14.0067</td>
</tr>
<tr>
<td>O</td>
<td>8</td>
<td>8</td>
<td>15.9994</td>
<td>15.9994</td>
</tr>
<tr>
<td>F</td>
<td>9</td>
<td>10</td>
<td>18.9984</td>
<td>18.9984</td>
</tr>
<tr>
<td>Ne</td>
<td>10</td>
<td>10</td>
<td>20.1797</td>
<td>20.1797</td>
</tr>
</tbody>
</table>

What does a mole look like?

![Image of beakers with different elements]

<table>
<thead>
<tr>
<th>Actual Mass and Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>55.85 g Fe</td>
</tr>
<tr>
<td>63.55 g Cu</td>
</tr>
<tr>
<td>118.69 g Sn</td>
</tr>
<tr>
<td>126.90 g I</td>
</tr>
</tbody>
</table>
04 Molar Mass

The molar mass of elements is found by looking at the atomic mass of the element on the periodic table.

If we wanted to know the molar mass of say magnesium or carbon we would discover that…

- **carbon** has a molar mass of 12.011g/mole
- **magnesium** has a molar mass of 24.305g/mole
Example 02

How many moles of carbon atoms are there in 24 g carbon atoms?

\[ 24 \text{ g C} \left( \frac{1 \text{ mol C}}{12 \text{ g C}} \right) = 2 \text{ mol C} \]

Example 03

How many atoms of carbon are there in 24 g carbon atoms?

\[ 24 \text{ g C} \left( \frac{1 \text{ mol C}}{12 \text{ g C}} \right) \left( \frac{6.022 \times 10^{23}}{1 \text{ mol}} \right) = 1.2 \times 10^{24} \text{ atoms C} \]
Lesson 02 and 03: Molar Mass and Unit Conversions and Multiple Unit Conversions

01 Molar Mass of an Element

The molar mass of elements is found by looking at the atomic mass of the element on the periodic table.
02 Molar Mass of a Compound

For any chemical compound that's not an element, we need to find the molar mass from the chemical formula. To do this, we need to remember that...

- molar masses of chemical compounds are equal to the sum of the molar masses of all the atoms in one molecule of that compound

<table>
<thead>
<tr>
<th>Example 01: Mole NaCl → Grams</th>
</tr>
</thead>
<tbody>
<tr>
<td>molar mass of NaCl = (1 mol Na) (\frac{23 \text{ g Na}}{1 \text{ mol Na}}) + (1 mol Cl) (\frac{35.5 \text{ g Cl}}{1 \text{ mol Cl}}) = 58.5 g/mol NaCl</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Example 02: FeCl₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>molar mass of FeCl₂ = (1 mol Fe) (\frac{56 \text{ g Fe}}{1 \text{ mol Fe}}) + (2 mol Cl) (\frac{35.5 \text{ g Cl}}{1 \text{ mol Cl}}) = 127 g/mol FeCl₂</td>
</tr>
</tbody>
</table>
Example 03: Zn(NO$_3$)$_2$

molar mass of Zn(NO$_3$)$_2$ = (1 mol Zn)$\left(\frac{65 \text{ g Zn}}{1 \text{ mol Zn}}\right)$ + 
(2 mol N)$\left(\frac{14 \text{ g N}}{1 \text{ mol N}}\right)$ + (6 mol O)$\left(\frac{16 \text{ g O}}{1 \text{ mol O}}\right)$

= 189 g/mol Zn(NO$_3$)$_2$
03 Unit Conversions

Moles to Mass and Mass to Moles

When converting from moles to mass or from mass to moles you need to use one of the following conversion factors...

\[
\frac{1 \text{ mole of substance}}{\text{molar mass of substance}} \quad \text{OR} \quad \frac{\text{molar mass of substance}}{1 \text{ mole of substance}}
\]

Determine the mass of 5.0 moles of oxygen gas (O₂).

\[
(5.0 \text{ mol O}_2) \left( \frac{32 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 160 \text{ g O}_2
\]

Determine the number of moles in 6.3 grams of H₂O.

\[
(6.3 \text{ g H}_2\text{O}) \left( \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g H}_2\text{O}} \right) = 0.35 \text{ mol H}_2\text{O}
\]
Number of Moles and Gas Volume

Remember that Avogadro's Hypothesis states that "equal volumes of gases at the same temperature and pressure contain the same number of molecules regardless of their chemical nature and physical properties".

1 mole of ANY gas at STP has a volume of 22.4 L

Conversion Factor

\[
\frac{1\text{ mol}}{22.4\text{ L}} \quad \text{OR} \quad \frac{22.4\text{ L}}{1\text{ mol}}
\]

We are going to use this information to help us solve the following two problems…
### What volume is occupied by 0.350 mol of carbon dioxide, CO₂, at STP?

\[
(0.350 \text{ mol CO}_2) \left( \frac{22.4 \text{ L}}{1 \text{ mol CO}_2} \right) = 7.84 \text{ L}
\]

### Calculate the number of moles for 5.00 ml of SO₂ gas.

\[
(5.00 \text{ ml SO}_2) \left( \frac{1 \text{ L}}{1000 \text{ ml}} \right) \left( \frac{1 \text{ mol SO}_2}{22.4 \text{ L SO}_2} \right) = 2.23 \times 10^{-4} \text{ mol SO}_2
\]
## Number of Moles and Number of Molecules or Atoms

### Conversion Factor

\[
\frac{1 \text{ mol molecules or atoms}}{6.022 \times 10^{23} \text{ molecules or atoms}} \quad \text{OR} \quad \frac{6.022 \times 10^{23} \text{ molecules or atoms}}{1 \text{ mol molecules or atoms}}
\]

### How many molecules are there in 0.5 moles of \( \text{H}_2\text{O} \)?

\[
(0.5 \text{ mol}) \left( \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \right) = 3 \times 10^{23} \text{ molecules}
\]

### How many atoms are there in 0.5 moles of \( \text{H}_2\text{O} \) molecules?

\[
(0.5 \text{ mol H}_2\text{O molecules}) \left( \frac{6.022 \times 10^{23} \text{ H}_2\text{O molecules}}{1 \text{ mol H}_2\text{O molecules}} \right) \left( \frac{3 \text{ H}_2\text{O atoms}}{1 \text{ H}_2\text{O molecule}} \right) = 5 \times 10^{24} \text{ H}_2\text{O atoms}
\]
04 Multiple Unit Conversions

Generally speaking, when converting the mass of a reactant to the mass of a product you must follow the procedure below…

When inter-converting between mass, volume and number of particles (molecules, atoms or ions) you must follow the procedure below…
Take a look at the following examples where we convert between the units in the previous diagram.

<table>
<thead>
<tr>
<th>You have 5.00 grams of calcium hydroxide, Ca(OH)$_2$. How many atoms do you have?</th>
</tr>
</thead>
</table>
| \[
(5.00 \text{ g Ca(OH)}_2) \left( \frac{1 \text{ mol Ca(OH)}_2}{74.1 \text{ g Ca(OH)}_2} \right) \left( \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} \right) = 4.06 \times 10^{22} \text{ atoms of Ca(OH)}_2
\] |

<table>
<thead>
<tr>
<th>What is the volume occupied by 5.0 grams of CO$_2$(g) at STP?</th>
</tr>
</thead>
</table>
| \[
(5.0 \text{ g CO}_2) \left( \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \right) \left( \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} \right) = 2.5 \text{ L CO}_2
\] |

<table>
<thead>
<tr>
<th>What is the volume occupied by 2.75 moles of methanol, CH$_3$OH(l)? (d = 0.7918 g/cm$^3$ = 0.7918 g/ml)</th>
</tr>
</thead>
</table>
| \[
(2.75 \text{ mol CH}_3\text{OH}) \left( \frac{32.0 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} \right) \left( \frac{1 \text{ ml CH}_3\text{OH}}{0.7918 \text{ g CH}_3\text{OH}} \right) = 111 \text{ ml CH}_3\text{OH}
\] |
### How many sulphur atoms are contained in 100.0 litres of SO₃(g) at STP?

\[
(100.0 \text{ L SO}_3) \left( \frac{1 \text{ mol SO}_3}{22.4 \text{ L SO}_3} \right) \left( \frac{6.022 \times 10^{23} \text{ molecules SO}_3}{1 \text{ mol SO}_3 \text{ molecules}} \right) \left( \frac{1 \text{ atom S}}{1 \text{ molecule SO}_3} \right) = 2.69 \times 10^{24} \text{ S atoms}
\]

### Al₂O₃(s) has a density of 3.97 g/ml. How many atoms of oxygen are in 100 milliliters of Al₂O₃(s)?

\[
(100 \text{ ml}) \left( \frac{3.97 \text{ g}}{\text{ ml}} \right) \left( \frac{1 \text{ mol Al}_2\text{O}_3}{102 \text{ g Al}_2\text{O}_3} \right) \left( \frac{6.022 \times 10^{23} \text{ molecules Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3 \text{ molecules}} \right) \left( \frac{3 \text{ atoms O}}{1 \text{ molecule Al}_2\text{O}_3} \right) = 7.03 \times 10^{24} \text{ O atoms}
\]
Lesson 04: Percent Composition

Percent composition is…
- percentage by mass of a component in a compound.

To calculate the percent composition of a component in a compound…
- Find the molar mass of the compound by adding up the masses of each atom in the compound using the periodic table.
- Find the mass of the component in the compound for which you are solving.
- Divide the mass of the component by the total molar mass of the compound and multiply by 100.

<table>
<thead>
<tr>
<th>Percent Composition Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>[ \text{percent composition} = \left( \frac{\text{mass of component}}{\text{total mass of compound}} \right) \times 100 ]</td>
</tr>
</tbody>
</table>
Example 01

What is the percentage composition of hydrogen and oxygen in $\text{H}_2\text{O}$?

**Hydrogen percent composition in $\text{H}_2\text{O}$…**

$$\% \text{ H} = \left( \frac{2 \times (1.01 \text{ g/mol H})}{18.02 \text{ g/mol } \text{H}_2\text{O}} \right) \times 100 = 11.2 \% \text{ H}$$

**Oxygen percent composition in $\text{H}_2\text{O}$…**

$$\% \text{ O} = \left( \frac{1 \times (16.00 \text{ g/mol O})}{18.02 \text{ g/mol } \text{H}_2\text{O}} \right) \times 100 = 88.79 \% \text{ O}$$
### Example 02

What is the percentage composition of each element in barium nitrate, \( \text{Ba(NO}_3\text{)}_2 \)?

<table>
<thead>
<tr>
<th>Element</th>
<th>Percentage Composition</th>
</tr>
</thead>
</table>
| **Barium** | \[
\% \text{ Ba} = \left( \frac{1 \times (137.33 \text{ g/mol Ba})}{261.35 \text{ g/mol Ba(NO}_3\text{)}_2} \right) \times 100 = 52.55 \% \text{ Ba}
\]
| **Nitrogen** | \[
\% \text{ N} = \left( \frac{2 \times (14.0 \text{ g/mol N})}{261.35 \text{ g/mol Ba(NO}_3\text{)}_2} \right) \times 100 = 10.72 \% \text{ N}
\]
| **Oxygen** | \[
\% \text{ O} = \left( \frac{6 \times (16.0 \text{ g/mol O})}{261.35 \text{ g/mol Ba(NO}_3\text{)}_2} \right) \times 100 = 36.73 \% \text{ O}
\]
Lesson 05: Empirical and Molecular Formulas

01 Types of Chemical Formulas

Chemical formulas can be divided into three types…
- **empirical formula**
- **molecular formula**
- **structural formula**

The first two are discussed in this lesson…
- **empirical formula**: shows the simplest whole number ratio for the elements in the compounds
- **molecular formula**: shows the correct ratio of elements in the compound (as it actually exists in nature)

<table>
<thead>
<tr>
<th>Compound</th>
<th>Molecular Formula</th>
<th>Empirical Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H₂O</td>
<td>H₂O</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C₂H₆O</td>
<td>CH₃O</td>
</tr>
<tr>
<td>Hydrogen Peroxide</td>
<td>H₂O₂</td>
<td>HO</td>
</tr>
<tr>
<td>Glucose</td>
<td>C₆H₁₂O₆</td>
<td>CH₂O</td>
</tr>
<tr>
<td>Methane</td>
<td>CH₄</td>
<td>CH₄</td>
</tr>
<tr>
<td>Ethane</td>
<td>C₂H₆</td>
<td>CH₃</td>
</tr>
</tbody>
</table>
The third type is discussed at the end of this course…

- **structural formula**: shows the correct ratio of elements in the compound and also displays the elements in their proper bonding arrangement in a diagrammatic fashion

<table>
<thead>
<tr>
<th>Compound</th>
<th>Molecular Formula</th>
<th>Structural Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methane</td>
<td>CH$_4$</td>
<td>H [\text{\quad H} \quad \text{\quad C} \quad \text{\quad H}] H [\text{\quad H}]</td>
</tr>
<tr>
<td>Propane</td>
<td>C$_3$H$_8$</td>
<td>H [\text{\quad H} \quad \text{\quad H} \quad \text{\quad C} \quad \text{\quad C} \quad \text{\quad C} \quad \text{\quad H}] H [\text{\quad H} \quad \text{\quad H} \quad \text{\quad H}]</td>
</tr>
</tbody>
</table>
02 Determining the Empirical and Molecular Formula of a Compound

In determining chemical formulae, assuming the molar mass of the compound is given, the…

- **FIRST STEP** is to determine the empirical formula
- **SECOND STEP** determined the molecular formula

If the molar mass is NOT given then it will be necessary to determine it too based on information such as…

- mass of given number of moles

\[ \text{molar mass} = \frac{\text{mass}}{\text{moles}} = \frac{x \ g}{x \ mol} = \frac{g}{\text{mol}} \]

- density of a gas at STP

\[ \text{molar mass} = (\text{density})(\text{molar volume}) = \left( x \ \frac{g}{L} \right) \left( \frac{22.4 \text{L}}{\text{mol}} \right) = \frac{g}{\text{mol}} \]

- mass and volume of a gas at STP

\[ \text{molar mass} = \frac{\text{mass}}{\text{volume}}(\text{molar volume}) = \left( x \ \frac{g}{L} \right) \left( \frac{22.4 \text{L}}{\text{mol}} \right) = \frac{g}{\text{mol}} \]
Example 01: Empirical Formula

Find the empirical formula for a compound containing 15.9% boron (B) and 84.1% fluorine (F).

- First make an assumption that we have 100 grams…
  
  100% → 100g  
  15.9% → 15.9g  
  84.1% → 84.1g

- Therefore we have…

\[
\left( 15.9 \text{ g B} \right) \left( \frac{1 \text{ mol B}}{10.8 \text{ g B}} \right) = 1.472 \text{ mol B} \\
\left( 84.1 \text{ g F} \right) \left( \frac{1 \text{ mol F}}{19.0 \text{ g F}} \right) = 4.268 \text{ mol F}
\]

- Determine the simplest whole number ratio by dividing by the smallest molar amount…

\[
B = \frac{1.472 \text{ mol}}{1.472 \text{ mol}} = 1.000 = 1 \\
F = \frac{4.268 \text{ mol}}{1.472 \text{ mol}} = 2.899 = 3
\]
Thus the empirical formula for this compound is…

\[ BF_3 \]

03 Empirical Formulae and Combustion Analysis

Combustion analysis is a technique used to determine the empirical formula of organic compounds, compounds that contain carbon and hydrogen.

When an organic compound is subject to combustion with oxygen in a special combustion apparatus all the…

- carbon is converted to \( CO_2 \), and the
- hydrogen to \( H_2O \)

The amount of…

- carbon produced can be determined by measuring the amount of \( CO_2 \) trapped by the sodium hydroxide, \( NaOH \)
• hydrogen produced by the amount of H₂O trapped by the magnesium perchlorate, Mg(ClO₄)₂

Example 02: Empirical Formula and Combustion Analysis

Consider the combustion of isopropyl alcohol. The sample is known to contain only carbon, hydrogen and oxygen. Combustion of 0.255 grams of isopropyl alcohol produces 0.561 grams of CO₂ and 0.306 grams of H₂O. Determine the empirical formula.

• first determine the amount of carbon present…

\[
(0.516 \text{ g CO}_2) \left( \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \right) \left( \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) \left( \frac{12.011 \text{ g C}}{1 \text{ mol C}} \right) = 0.154 \text{ g C}
\]

• then determine the amount of hydrogen present…

\[
(0.306 \text{ g H}_2\text{O}) \left( \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \right) \left( \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) \left( \frac{1.0079 \text{ g H}}{1 \text{ mol H}} \right) = 0.0342 \text{ g H}
\]

• then determine the missing mass…
(given mass) – (mass of C and H) = mass of O

\[(0.255 \text{ g}) – (0.154 \text{ g } C + 0.0342 \text{ g } H) = 0.067 \text{ g } O\]

- then determine the moles of each atom present…

\[
\begin{align*}
(0.154 \text{ g } C) \left(\frac{1 \text{ mol } C}{12.011 \text{ g } C}\right) &= 0.0128 \text{ moles } C \\
(0.034 \text{ g } H) \left(\frac{1 \text{ mol } H}{1.0079 \text{ g } H}\right) &= 0.034 \text{ moles } H \\
(0.067 \text{ g } O) \left(\frac{1 \text{ mol } O}{15.9994 \text{ g } O}\right) &= 0.0042 \text{ moles } O
\end{align*}
\]

- convert fractions to whole numbers…

\[
\begin{align*}
C &= \frac{0.0128 \text{ moles}}{0.0042 \text{ moles}} = 3.0 \\
H &= \frac{0.034 \text{ moles}}{0.0042 \text{ moles}} = 8.1 \\
O &= \frac{0.0042 \text{ moles}}{0.0042 \text{ moles}} = 1.0
\end{align*}
\]

- and finally the empirical formula is…

\[C_3\text{H}_8\text{O}\]
Example 03: Molecular Formula (Molar Mass NOT Given)

A gas has an empirical formula of POF$_3$. If 0.350 L of the gas at STP has a mass 1.62 g, what is the molecular formula of the compound?

- First determine the empirical mass of POF$_3$...

  $104.0 \text{ g}$

- Second, determine the molar mass...

  $\text{molar mass} = (\text{density})(\text{molar volume})$

  $\text{molar mass} = \left( \frac{1.62 \text{ g}}{0.350 \text{ L}} \right) \left( \frac{22.4 \text{ L}}{1 \text{ mol}} \right) = 104 \text{ g/mol}$

- Finally determine how much larger the molecular mass is compared to the empirical mass. This is the number you will multiply all subscripts by.
\[
\frac{\text{molar mass}}{\text{empirical mass}} = \frac{104 \text{ g/mol}}{104 \text{ g/mol}} = 1
\]

- In this case the molecular formula is the same as the empirical formula and therefore there is no need to change any subscripts.

\[
P_OF_3
\]

---

**Example 04: Molecular Formula (Molar Mass Given)**

Vitamin C (ascorbic acid) contains 40.92 % C, 4.58 % H, and 54.50 % O, by mass. The experimentally determined molar mass is 176 g/mol. What is the molecular formula for ascorbic acid?

- First make an assumption that we have 100 grams.

- Therefore, we have...

\[
\begin{align*}
(40.92 \text{ g C}) \left( \frac{1 \text{ mol C}}{12.011 \text{ g C}} \right) &= 3.407 \text{ mol C} \\
(4.58 \text{ g H}) \left( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) &= 4.544 \text{ mol H} \\
(54.50 \text{ g O}) \left( \frac{1 \text{ mol O}}{15.9994 \text{ g O}} \right) &= 3.406 \text{ mol O}
\end{align*}
\]

- Determine the simplest whole number ratio by dividing by the smallest molar amount...
• Since we cannot have fractional amounts of atoms in a compound, we need to multiply the relative amounts of each atom by a number which gets rid of the fractional amounts. In this case multiply by three…

\[
\begin{align*}
C &= (1.001)(3) = 3 \\
H &= (1.333)(3) = 4 \\
O &= (1.000)(3) = 3
\end{align*}
\]

• Thus the empirical formula for ascorbic acid is…

\[
C_3H_4O_3
\]

• But what about the molecular formula?

• We are told that the experimentally determined molecular mass is 176 g/mol. What is the molecular mass of our empirical formula? The molecular mass from our empirical formula is…
\[ (3) \left( \frac{12.011 \text{ g C}}{1 \text{ mol C}} \right) + (4) \left( \frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) + \\
(3) \left( \frac{15.9994 \text{ g O}}{1 \text{ mol O}} \right) = 88.062 \text{ g/mol C}_3\text{H}_4\text{O}_3 \]

- The molecular mass from our empirical formula is significantly lower than the experimentally determined value…

\[
\text{176 g/mol versus 88.062 g/mol}
\]

- What is the ratio between the two values?...

\[
\frac{176 \text{ g/mol}}{88.062 \text{ g/mol}} = 2.00
\]

- Thus the molecular formula for ascorbic acid is…

\[
2(C_3H_4O_3) = C_6H_8O_6
\]
Solutions are homogeneous mixtures of two or more substances. Usually the solution is a liquid, but this is not always the case.

Normally, a mixture has two components…
- **solute**: the substance you are dissolving
- **solvent**: the substance you are dissolving into
Since solutions are a mixture, and therefore the ratio of solute to solvent varies. This **ratio of solute to solvent** is called **concentration**.
2. Making a Solution

A **volumetric flask** is used to make up solutions of fixed volume very accurately.

To make up a solution…
- first add your solid to a flask with an appropriate volume
- dissolve the solid completely in a little water by shaking
- carefully fill the flask to just below the mark
• add distilled water a drop at a time until the bottom of the meniscus lines up exactly with the mark on the neck of the flask

• after the final dilution, remember to mix your solution thoroughly, by inverting the flask and shaking.
3. Qualitative Description of Concentration

To concentrate a solution, one must either…
- add more solute (most common), or
- reduce the amount of solvent

To dilute a solution, one must either…
- add more solvent (most common) or
- reduce the amount of solute

Example of Dilute and Concentrated Solutions
4. Quantitative Description of Concentration

For scientific or technical applications, a qualitative account of concentration is almost never sufficient; therefore quantitative measures are needed to describe concentration.

There are a number of ways to express concentration. Which one we choose often depends on convenience.

In chemistry 11 and 12 we will deal exclusively with molarity…

<table>
<thead>
<tr>
<th>Some Common Concentration Measures</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Molarity</td>
</tr>
<tr>
<td>moles of solute</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>L of solution</td>
</tr>
<tr>
<td>2. Molality</td>
</tr>
<tr>
<td>moles of solute</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>kg of solvent</td>
</tr>
<tr>
<td>3. Parts per million</td>
</tr>
<tr>
<td>mass of solute</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>mass of solution • 10^6</td>
</tr>
<tr>
<td>4. Parts per billion</td>
</tr>
<tr>
<td>mass of solute</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>mass of solution • 10^9</td>
</tr>
<tr>
<td>5. Parts per trillion</td>
</tr>
<tr>
<td>mass of solute</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>mass of solution • 10^{12}</td>
</tr>
<tr>
<td>6. Mass percent</td>
</tr>
<tr>
<td>mass of solute</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>mass of solution • 100</td>
</tr>
<tr>
<td>7. Volume percent</td>
</tr>
<tr>
<td>mL of solute</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>mL of solution • 100</td>
</tr>
</tbody>
</table>
5. Molarity Calculations

Example 01

What is the molarity of a solution made by dissolving 20 grams of NaCl in 100 ml’s of water?

\[
20 \text{ grams NaCl} \left( \frac{1 \text{ mol NaCl}}{58.5 \text{ grams NaCl}} \right) \left( \frac{1}{0.1 \text{ L}} \right) = 3.4 \text{ moles/L NaCl or 3.4 M NaCl}
\]

Example 02

Calculate the number of moles of CaCl₂ in 0.78 litres of a 3.5 M solution…

\[
0.78 \text{ L} \left( \frac{3.5 \text{ moles CaCl}_2}{1 \text{ L}} \right) = 2.73 \text{ moles CaCl}_2
\]

Example 03

How many litres of a 2.0 M solution of HNO₃ do we need to have 5 moles of HNO₃?

\[
(5 \text{ moles}) \left( \frac{1 \text{ L}}{2 \text{ moles}} \right) = 2.5 \text{ L HNO}_3
\]
6. Dilution

For convenience, solutions are either purchased or prepared in concentrated **stock solutions** which must be diluted prior to use.

When we take a **sample** of a stock solution we have a certain number of moles of molecules in that sample. Dilution alters the molarity or concentration of the solution but not the total number of moles of molecules in the solution.

One of the standard equations for determining the effects of dilution upon a sample is to set up an equation comparing concentration and volume before and after dilution. Since concentration times volume gives us the total number of moles in the sample, and since this does not change, this value before and after dilution are equal…

\[
M_{\text{before}} V_{\text{before}} = M_{\text{after}} V_{\text{after}}
\]

\[
\text{moles}_{\text{before}} = \text{moles}_{\text{after}}
\]
Example 04

How much of a 5 M stock solution of NaCl will you need to make up 250 ml of a 1.5 M solution?

\[ M_{\text{before}} V_{\text{before}} = M_{\text{after}} V_{\text{after}} \]

\[ V_{\text{before}} = \frac{M_{\text{after}} V_{\text{after}}}{M_{\text{before}}} = \frac{(1.5 \text{ M})(0.25 \text{ L})}{(5 \text{ M})} = 0.075 \text{ L} \]

Thus, we would need 0.075 litres of our 5 M NaCl stock solution. The rest of the 0.25 L volume is made up by the addition of water…

\[ 0.25 \text{ L} - 0.075 \text{ L} = 0.175 \text{ L} \]

So we would take 0.075 litres of stock 5M NaCl solution and add that to 0.175 litres of water for a final volume of 0.25 litres with a final concentration of 1.5 M.