Chapter 3

Stoichiometry of Formulas and Equations
Chapter 3 Outline:
Mole - Mass Relationships in Chemical Systems

3.1 The Mole

3.2 Determining the Formula of an Unknown Compound

3.3 Writing and Balancing Chemical Equations

3.4 Calculating Amounts of Reactant and Product

3.5 Fundamentals of Solution Stoichiometry
Atoms are so small that a convenient way of counting numbers of atoms is necessary. 

*Best: count by weight!*

**Mole (mol):** number of objects (atoms, molecules, etc.) equal to the number of atoms present in *exactly* 12 grams of carbon-12.

\[
\text{Mass of C-12 atom} = 1.9926 \times 10^{-23} \text{ g}
\]

1 mol of C-12 atoms = 12 g of C-12 atoms = \(6.0221 \times 10^{23}\) C-12 atoms

*This is known as Avogadro’s Number*

\[
N_A \text{ (or } N_0) = 6.022 \times 10^{23} \text{ items/mol}
\]
A mole of various elements

Each contains $6.022 \times 10^{23}$ atoms!
One mole of some familiar substances.

CaCO$_3$
100.09 g

Water
18.02 g

Oxygen
32.00 g

Copper
63.55 g
Amount of substance in mol: \( n \)
Number of substance ‘pieces’ \( N = nN_A \)
And \( n = \frac{N}{N_A} \)

So \( 1.27 \times 10^{19} \) atoms of Na = \( 2.11 \times 10^{-5} \) mol of Na

But can we relate mass to moles?
YES!

*If we know the weight of one unit of what we are interested in we can multiplying that by \( N_A \) will give us the weight of one mole of it!*

This is the MOLAR MASS
Molar Mass

The SUM of the ATOMIC masses of the elements that make up that molecule

\[
\text{Al}_2(\text{SO}_4)_3
\]

- \(2 \text{ Al} \times 26.98 \text{ g/mol} = 53.96 \text{ g Al}\)
- \(3 \text{ S} \times 32.06 \text{ g/mol} = 96.18 \text{ g S}\)
- \(12 \text{ O} \times 16.00 \text{ g/mol} = 192.00 \text{ g O}\)

\[\text{MM} = 342.14 \text{ g/mol}\]
# Molar Mass and Chemical Formulas

## Information Contained in the Chemical Formula of Glucose \( \text{C}_6\text{H}_{12}\text{O}_6 \) 

\( MM = 180.16 \text{ g/mol} \)

<table>
<thead>
<tr>
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<th>Carbon (C)</th>
<th>Hydrogen (H)</th>
<th>Oxygen (O)</th>
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</thead>
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<td>Atoms/molecule of compound</td>
<td>6 atoms</td>
<td>12 atoms</td>
<td>6 atoms</td>
</tr>
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<td>Moles of atoms/mole of compound</td>
<td>6 mol of atoms</td>
<td>12 mol of atoms</td>
<td>6 mol of atoms</td>
</tr>
<tr>
<td>Atoms/mole of compound</td>
<td>(6(6.022 \times 10^{23})) atoms</td>
<td>(12(6.022 \times 10^{23})) atoms</td>
<td>(6(6.022 \times 10^{23})) atoms</td>
</tr>
<tr>
<td>Mass/molecule of compound</td>
<td>(6(12.01 \text{ amu}) = 72.06 \text{ amu})</td>
<td>(12(1.008 \text{ amu}) = 12.10 \text{ amu})</td>
<td>(6(16.00 \text{ amu}) = 96.00 \text{ amu})</td>
</tr>
<tr>
<td>Mass/mole of compound</td>
<td>72.06 g</td>
<td>12.10 g</td>
<td>96.00 g</td>
</tr>
</tbody>
</table>
Summary of the mass-mole-number relationships for compounds.
Central Concept:

Almost ALL chemical calculations will rest on this interconversion.
Terms, Terms, Terms…

• Atomic weight/Atomic mass
• Molecular weight/Molecular mass
• Formula Weight/Formula mass

Bottom Line: They all mean the same thing!
Mass Percent

Since we can measure mass directly and we now know how to convert that to moles it is common practice to refer to the mass composition of a compound when experimentally determining the identity of an unknown. So, Mass Percent from a chemical formula is simply done on a per element basis as:

\[
\text{Mass % of element } X = \frac{\text{atoms of } X \text{ in formula} \times \text{atomic mass of } X \text{ (amu)}}{\text{molecular (or formula) mass of compound (amu)}} \times 100
\]
Mass Percent: Aluminum Sulfate

For $\text{Al}_2(\text{SO}_4)_3$:

$\text{Mass of } \text{Al} = 2 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 53.96 \text{ g Al}$

$\text{Mass of } \text{S} = 3 \text{ mol S} \times \frac{32.06 \text{ g S}}{1 \text{ mol S}} = 96.18 \text{ g S}$

$\text{Mass of } \text{O} = 12 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 192.00 \text{ g O}$

So:

$\% \text{Al} = \frac{53.96 \text{ g Al}}{342.14 \text{ g } \text{Al}_2(\text{SO}_4)_3} \times 100 = \boxed{15.77\% \text{ Al}}$

$\% \text{S} = \frac{96.18 \text{ g S}}{342.14 \text{ g } \text{Al}_2(\text{SO}_4)_3} \times 100 = \boxed{28.11\% \text{ S}}$

$\% \text{O} = \frac{192.00 \text{ g O}}{342.14 \text{ g } \text{Al}_2(\text{SO}_4)_3} \times 100 = \boxed{56.12\% \text{ O}}$
Empirical and Molecular Formulas

**Empirical Formula -**

The simplest formula for a compound that agrees with the elemental analysis and gives rise to the *smallest set of whole numbers of atoms (lowest common multiple).*

**Molecular Formula -**

The formula of the compound as it exists; it *may* be a multiple of the empirical formula.

*Generally: for Type I and II compounds, the empirical and molecular formulas are the same; Type III compounds is where we usually see a difference.*

\[
\text{n Empirical Formula Units } A_x B_y C_z = \text{Molecular Formula } A_{nx} B_{ny} C_{nz}
\]

\[
\text{n} = \frac{\text{MM of the actual compound}}{\text{MM of the EF}}
\]
Determining an Empirical Formula from Masses of Elements

Elemental analysis of a sample of an ionic compound showed 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What is the empirical formula and name of the compound?

Strategy: To be able to get to the EF we need the molar ratio of atoms. This can be determined from the amounts of the elements present or the compound’s percent composition (later). Once we find the relative number of moles of each element, we can divide by the lowest mol amount to find the relative mol ratios (empirical formula).
Determining an Empirical Formula from Masses of Elements: Solution

\[ 2.82 \text{ g Na} \times \frac{\text{mol Na}}{22.99 \text{ g Na}} = 0.123 \text{ mol Na} \]

\[ 4.35 \text{ g Cl} \times \frac{\text{mol Cl}}{35.45 \text{ g Cl}} = 0.123 \text{ mol Cl} \]

\[ 7.83 \text{ g O} \times \frac{\text{mol O}}{16.00 \text{ g O}} = 0.489 \text{ mol O} \]

\[
\text{Na:Cl:O ratio} = \frac{0.123 \text{ mol}}{0.123} : \frac{0.123 \text{ mol}}{0.123} : \frac{0.489 \text{ mol}}{0.123}
\]

\[ \text{Na}_{1.00} \text{Cl}_{1.00} \text{O}_{3.98} \rightarrow \text{NaClO}_4 \]

NaClO\(_4\) is sodium perchlorate.
Determining a Molecular Formula from Elemental Analysis (% Composition) and Molar Mass

During physical activity, lactic acid (\(\text{MM} = 90.08 \text{ g/mol}\)) forms in muscle tissue and is responsible for muscle soreness. Elemental analysis shows that this compound contains 40.0 mass % C, 6.71 mass % H, and 53.3 mass % O.

(a) Determine the empirical formula of lactic acid.

(b) Determine the molecular formula.

Strategy: The strategy here is to remember that percent mean out of 100, so if we assume we start with 100 g of lactic acid we can convert the percentages directly into the mass of each element. From there we convert mass of each to moles, get a ratio and convert to integer subscripts.

To get the molecular formula we will divide molar mass by the empirical formula mass to get the multiplier \(n\), which will then give us the molecular formula.
Determining a Molecular Formula from Elemental Analysis and Molar Mass: Solution

Based on the percentages, in 100.0 g of lactic acid, there are: 40.0 g C, 6.71 g H, and 53.3 g O

Convert the grams to moles and get a ratio for the empirical formula.

\[
\begin{align*}
40.0 \text{ g C} \times \frac{\text{mol C}}{12.01 \text{ g C}} &= 3.33 \text{ mol C} \\
6.71 \text{ g H} \times \frac{\text{mol H}}{1.008 \text{ g H}} &= 6.66 \text{ mol H} \\
53.3 \text{ g O} \times \frac{\text{mol O}}{16.00 \text{ g O}} &= 3.33 \text{ mol O}
\end{align*}
\]

\[
\frac{3.33 \text{ mol C}}{3.33} : \frac{6.66 \text{ mol H}}{3.33} : \frac{3.33 \text{ mol O}}{3.33}
\]

So: \( \text{C:H:O} = \text{empirical formula} \)

\[
n = \frac{\text{molar mass of lactate}}{\text{molar mass of CH}_2\text{O}} = \frac{90.08 \text{ g}}{30.03 \text{ g}} = 3
\]

So: \( \text{C}_3\text{H}_6\text{O}_3 \) is the molecular formula

The empirical formulas of organic molecules (which contain C, H, and O primarily) are usually determined by combustion analysis. A sample is burned (combusted) in a stream of oxygen gas; all elements present are converted to their oxides, most notably carbon to CO$_2$ and H to H$_2$O (other elements present are generally determined other ways). IF oxygen is present in the original sample it must be determined by mass difference (since an excess of oxygen was used in the combustion)

\[ C_x H_y E_z + \left(x + \frac{y}{2}\right)O_2 \rightarrow x\ CO_2(g) + \frac{y}{2}\ H_2O(g) \]

E is determined by difference or other method

Chemical Quantities and Stoichiometry
Determining a Molecular Formula from Combustion Analysis

Vitamin C ($M = 176.12 \text{ g/mol}$) is a compound of C, H, and O found in many natural sources, especially citrus fruits. When a 1.000-g sample of vitamin C is placed in a combustion chamber and burned, 1.500 g of carbon dioxide and 410. mg of water were obtained. What is the molecular formula of vitamin C?

**Strategy:** We need to find out how much of the sample mass is due to each element present; this requires converting the mass of both the CO$_2$ and H$_2$O to moles of each, from moles of each to moles of C and H, and from that to mass of C and H. The difference in mass between these and the original sample will be the amount of oxygen present; converting that to moles will allow us to set up our empirical formula ratio and the empirical formula. Using the actual molar mass and empirical formula mass, we will determine the molecular formula.
Determining a Molecular Formula from Combustion Analysis:
Solution

\[
1.50 \text{ g } \text{CO}_2 \times \frac{1 \text{ mol } \text{CO}_2}{44.01 \text{ g } \text{CO}_2} \times \frac{1 \text{ mol } \text{C}}{1 \text{ mol } \text{CO}_2} = \frac{0.0341 \text{ mol } \text{C}}{1 \text{ mol } \text{C}} \times \frac{12.01 \text{ g } \text{C}}{1 \text{ mol } \text{C}} = 0.409 \text{ g } \text{C}
\]

\[
0.410 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} \times \frac{2 \text{ mol } \text{H}}{1 \text{ mol } \text{H}_2\text{O}} = \frac{0.0456 \text{ mol } \text{H}}{1 \text{ mol } \text{H}} \times \frac{1.008 \text{ g } \text{H}}{1 \text{ mol } \text{H}} = 0.046 \text{ g } \text{H}
\]

\(\text{O Is determined by difference: } 1.000 \text{ g sample } - 0.409 \text{ g C } - 0.046 \text{ g H } = 0.545 \text{ g O}\)

\[
0.545 \text{ g } \text{O} \times \frac{1 \text{ mol } \text{O}}{16.00 \text{ g } \text{O}} = \frac{0.0341 \text{ mol } \text{O}}{1 \text{ mol } \text{O}}
\]
Determining a Molecular Formula from Combustion Analysis: Solution (cont’d)

\[
\text{C:H:O ratio} = \frac{\text{0.0341 mol C}}{\text{0.0341}} : \frac{\text{0.0456 mol H}}{\text{0.0341}} : \frac{\text{0.0341 mol O}}{\text{0.0341}}
\]

C:H:O ratio = 1 : 1.3 : 1; multiply each by 3 to get whole numbers

\[\text{Empirical Formula} = C_3 H_4 O_3 \, (\text{MM} = 88.06 \, \text{g/mol}) \]

\[
n = \frac{\text{Actual MM}}{\text{MM of EF}} = \frac{176.12 \, \text{g/mol}}{88.06 \, \text{g/mol}} = 2
\]

\[\text{Molecular Formula} = C_6 H_8 O_6 \]
Balancing Chemical Equations

Translate the statement and write correct chemical formulas

Balance the atoms; start with a metal if possible; treat polyatomics that don’t change as if they were single elements

Adjust the coefficients; use fractions if necessary but eliminate at end

Check the atom balance

Specify states of matter
Balancing Chemical Equations

PROBLEM: Within the cylinders of a car’s engine, the hydrocarbon octane (C\textsubscript{8}H\textsubscript{18}), one of many components of gasoline, mixes with oxygen from the air and burns to form carbon dioxide and water vapor. Write a balanced equation for this reaction.

SOLUTION:

\[ \text{C}_8\text{H}_{18} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]

\[ \text{C}_8\text{H}_{18} + \frac{25}{2} \text{O}_2 \rightarrow 8 \text{CO}_2 + 9\text{H}_2\text{O} \]

\[ 2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O} \]

\[ 2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O} \]

\[ 2\text{C}_8\text{H}_{18}(l) + 25\text{O}_2(g) \rightarrow 16\text{CO}_2(g) + 18\text{H}_2\text{O}(g) \]
Stoichiometry

From two Greek words: *stoicheion* ("element") and *metron* ("measure").

Richter: *Stoichiometry is the science of measuring the quantitative proportions or mass ratios in which chemical elements stand to one another.*

Stoichiometry deals with calculations about the masses, moles, and volumes of reactants and products involved in a chemical reaction. It is the ‘accounting’ of chemistry.

Most stoichiometric problems present you with a certain amount of a reactant (or reactants) and then ask how much of a product can be formed, or present you with a desired quantity of product to make and ask you how much reactant you need to achieve this goal.
Chemical Quantity Interconversions

Atoms or molecules A \( \leftrightarrow \) Vol Soln A (L) \( \leftrightarrow \) Vol Soln B (L) \( \leftrightarrow \) Atoms or molecules B

Vol Gas A (L) \( \leftrightarrow \) moles A \( \leftrightarrow \) moles B \( \leftrightarrow \) Vol Gas B (L)

Vol Gas or Liq A (mL) \( \leftrightarrow \) grams A \( \leftrightarrow \) grams B \( \leftrightarrow \) Vol Gas or Liq B (mL)

\[ MM = \text{molar mass (g/mol)} \]
\[ M = \text{molarity (mol/L)} \]
\[ V_m = \text{Molar volume (L/mol)} \]
\[ d = \text{density (g/mL)} \]
\[ N_A = \text{Avogadro's number (atoms/mol)} \]
Summary of the Mass-Mole-Number Relationships In A Chemical Reaction.
The "Stoichiometric Ratio"

3H₂(g) + N₂(g) → 2NH₃(g)
Stoichiometric Calculations

1. Write and balance your chemical equation;
2. Calculate molar masses of any materials for which have been given amounts OR will have determine amounts for;
3. Convert any masses given to moles;
4. Use balanced equation to determine stoichiometric ratios needed;
5. Calculate moles of desired materials;
6. Calculate grams of desired materials.
Stoichiometric Calculation Example

If a solution containing 14.7 g of silver nitrate is treated with a solution of sodium phosphate, how much silver phosphate will be made?

\[
\text{Na}_3\text{PO}_4(\text{aq}) + 3\text{AgNO}_3(\text{aq}) \rightarrow \text{Ag}_3\text{PO}_4(\text{s}) + 3\text{NaNO}_3(\text{aq})
\]

<table>
<thead>
<tr>
<th>MM</th>
<th>168.9 g/mol</th>
<th>415.7 g/mol</th>
</tr>
</thead>
<tbody>
<tr>
<td>(g) (start)</td>
<td>excess</td>
<td>14.7 g</td>
</tr>
<tr>
<td>(mol) (start)</td>
<td>excess</td>
<td>0.0870 mol</td>
</tr>
<tr>
<td>(mol) (end)</td>
<td>excess**</td>
<td>0 mol</td>
</tr>
<tr>
<td>(g) (start)</td>
<td>excess**</td>
<td>0 g</td>
</tr>
</tbody>
</table>
\[ \text{Na}_3\text{PO}_4(\text{aq}) + 3\text{AgNO}_3(\text{aq}) \rightarrow \text{Ag}_3\text{PO}_4(\text{s}) + 3\text{NaNO}_3(\text{aq}) \]

\[
14.7 \text{ g } \text{AgNO}_3 \times \frac{1 \text{ mole } \text{AgNO}_3}{168.9 \text{ g } \text{AgNO}_3} = 0.0870 \text{ mol } \text{AgNO}_3
\]

\[
0.0870 \text{ mol } \text{AgNO}_3 \times \frac{1 \text{ mole } \text{Ag}_3\text{PO}_4}{3 \text{ mol } \text{AgNO}_3} = 0.0290 \text{ mol } \text{Ag}_3\text{PO}_4
\]

\[
0.0290 \text{ mol } \text{Ag}_3\text{PO}_4 \times \frac{415.7 \text{ g } \text{Ag}_3\text{PO}_4}{1 \text{ mol } \text{Ag}_3\text{PO}_4} = 12.1 \text{ g } \text{Ag}_3\text{PO}_4
\]

\[ \text{g } \text{AgNO}_3 \rightarrow \text{mol } \text{AgNO}_3 \rightarrow \text{mol } \text{Ag}_3\text{PO}_4 \rightarrow \text{g } \text{Ag}_3\text{PO}_4 \]
**Limiting Reagent**

Typically, a chemical reaction is carried out such that one of the starting materials is present in a smaller stoichiometric amount than the others; this is said to be the *LIMITING REAGENT*.

- So-named because it LIMITS the amount of product(s) you can form;
- Based on both molar amounts *given* and stoichiometric amounts *needed*;
- Limiting reagent problems are easy to spot: you will be given amounts of more than one of the starting materials
- **Caution**: the limiting reagent must be determined by calculation!
Limiting Reagent Calculations

• Carry out steps 1-3 as outlined for standard stoichiometric calculations;

• Determine the limiting reagent using either: (i) the “Need vs. Have” method; (ii) the “Product” method; or (iii) the “Ratio” method (not discussed and NOT recommended);

• Calculate moles of desired materials based on your limiting reagent;

• Calculate grams of desired materials based on your limiting reagent;
“Need vs. Have” Method

- Based on knowing how much of a given starting material you have and comparing it with how much you need;
- Using any one of your starting materials (SM #1), calculate how much of the other starting material (SM #2) you will NEED to completely react with it; be sure to use stoichiometric ratios from your balanced equation!
- Compare the amount NEEDED for SM #2 (i.e. what you just calculated) with the actual amount of SM #2 you HAVE;
- If NEED > HAVE, SM #2 is your limiting reagent;
- If HAVE > NEED, SM #1 is your limiting reagent;
- Finish all stoichiometric calculations based on your limiting reagent
“Product” Method

- Based on the fact that you can make no more of a given product than your limiting reagent allows;
- Start by picking any one of the products expected in the reaction;
- Using the first starting material (SM #1), calculate how much of the product you chose it will make; repeat the same calculation using the other starting material (SM #2).
- The starting material that gives you the SMALLEST amount of product is the limiting reagent;

**NOTE:** this amount is the MAXIMUM amount of that product you can possibly make (the theoretical yield)
# Limiting Reagent Example

If a solution containing 14.7 g of silver nitrate is treated with a solution of 21.4 g of sodium phosphate, how much silver phosphate will be made? How much of the “reagent in excess” (RIE) will remain at reactions end?

\[
\text{Na}_3\text{PO}_4(\text{aq}) + 3\text{AgNO}_3(\text{aq}) \rightarrow \text{Ag}_3\text{PO}_4(\text{s}) + 3\text{NaNO}_3(\text{aq})
\]

<table>
<thead>
<tr>
<th></th>
<th>Na(_3)PO(_4) (aq)</th>
<th>3AgNO(_3) (aq)</th>
<th>Ag(_3)PO(_4) (s)</th>
<th>3NaNO(_3) (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(g) (start)</td>
<td>21.4 g</td>
<td>14.7 g</td>
<td>0 g</td>
<td></td>
</tr>
<tr>
<td>(\text{mol (start)})</td>
<td>0.131 mol</td>
<td>0.0870 mol</td>
<td>0 mol</td>
<td></td>
</tr>
<tr>
<td>(\text{mol (end)})</td>
<td>??</td>
<td>??</td>
<td>?? mol</td>
<td></td>
</tr>
<tr>
<td>(g) (end)</td>
<td>??</td>
<td>??</td>
<td>?? g</td>
<td></td>
</tr>
</tbody>
</table>

ONE of the starting materials **WILL** be **COMPLETELY** used up (the limiting reagent)
Limiting Reagent: “Need vs. Have”

How much $\text{Na}_3\text{PO}_4$ will I need to react with all the $\text{AgNO}_3$ I have?

$\text{Na}_3\text{PO}_4(\text{aq}) + 3\text{AgNO}_3(\text{aq}) \rightarrow \text{Ag}_3\text{PO}_4(\text{s}) + 3\text{NaNO}_3(\text{aq})$

$$14.7 \text{ g } \text{AgNO}_3 \times \frac{1 \text{ mole } \text{AgNO}_3}{168.9 \text{ g } \text{AgNO}_3} = 0.0870 \text{ mol } \text{AgNO}_3$$

$$0.0870 \text{ mol } \text{AgNO}_3 \times \frac{1 \text{ mole } \text{Na}_3\text{PO}_4}{3 \text{ mol } \text{AgNO}_3} = 0.0290 \text{ mol } \text{Na}_3\text{PO}_4 \text{ NEEDED}$$

I HAVE 0.131 mol of $\text{Na}_3\text{PO}_4$ but only NEED 0.0290 mol; since I HAVE more $\text{Na}_3\text{PO}_4$ than I NEED it is NOT the limiting reagent; therefore AgNO$_3$ IS the limiting reagent!
**Limiting Reagent: “Need vs. Have”**

Now, calculate the amount of desired product based on your limiting reagent (AgNO₃):

\[
0.0870 \text{ mol AgNO}_3 \times \frac{1 \text{ mole Ag}_3\text{PO}_4}{3 \text{ mol AgNO}_3} = 0.0290 \text{ mol Ag}_3\text{PO}_4
\]

\[
0.0290 \text{ mol Ag}_3\text{PO}_4 \times \frac{415.7 \text{ g Ag}_3\text{PO}_4}{1 \text{ mol Ag}_3\text{PO}_4} = 12.1 \text{ g Ag}_3\text{PO}_4
\]

Therefore, the MAXIMUM amount of Ag₃PO₄ you will be able to make is 12.1 g; this is the *theoretical yield* for this reaction.
**Limiting Reagent: “Product” Method**

How much $\text{Ag}_3\text{PO}_4$ can I make with the amount of $\text{Na}_3\text{PO}_4$ I have to start with?

How much $\text{Ag}_3\text{PO}_4$ can I make with the amount of $\text{AgNO}_3$ I have to start with?

Based on $\text{Na}_3\text{PO}_4$ to start:

$$0.131 \text{ mol } \text{Na}_3\text{PO}_4 \times \frac{1 \text{ mol } \text{Ag}_3\text{PO}_4}{1 \text{ mol } \text{Na}_3\text{PO}_4} = 0.131 \text{ mol } \text{Ag}_3\text{PO}_4 \text{ can be made}$$

Based on $\text{AgNO}_3$ to start:

$$0.0870 \text{ mol } \text{AgNO}_3 \times \frac{1 \text{ mole } \text{Ag}_3\text{PO}_4}{3 \text{ mol } \text{AgNO}_3} = 0.0290 \text{ mol } \text{Ag}_3\text{PO}_4 \text{ can be made}$$

Since I will make LESS product from the starting amount of $\text{AgNO}_3$ than from the starting amount of $\text{Na}_3\text{PO}_4$, $\text{AgNO}_3$ IS the limiting reagent!
Limiting Reagent: “Product” Method

As before, calculate the amount of desired product based on your limiting reagent (AgNO₃):

\[ 0.0870 \text{ mol AgNO}_3 \times \frac{1 \text{ mole Ag}_3\text{PO}_4}{3 \text{ mol AgNO}_3} = 0.0290 \text{ mol Ag}_3\text{PO}_4 \]

\[ 0.0290 \text{ mol Ag}_3\text{PO}_4 \times \frac{415.7 \text{ g Ag}_3\text{PO}_4}{1 \text{ mol Ag}_3\text{PO}_4} = 12.1 \text{ g Ag}_3\text{PO}_4 \]

Note that the results obtained are identical for the two methods!
“Reagent in Excess” Determination

Quite often you will be asked to determine how much of the excess reagent will be left at reactions end; this is simply done by determining how much of the RIE was needed to react with the limiting reagent provided and taking the difference:

0.0870 mol AgNO₃ will require:

\[ 0.0870 \text{ mol AgNO}_3 \times \frac{1 \text{ mole Na}_3\text{PO}_4}{3 \text{ mol AgNO}_3} = 0.0290 \text{ mol Na}_3\text{PO}_4; \]

Amt of Na₃PO₄ to start = 0.131 mol

\[ \text{Amt of Na}_3\text{PO}_4 \text{ left} = 0.131 \text{ mol} - 0.0290 \text{ mol} = 0.102 \text{ mol Na}_3\text{PO}_4 \]

\[ \text{Amt of Na}_3\text{PO}_4 \text{ left} = 16.7 \text{ g Na}_3\text{PO}_4 \]
The Effect Of Side Reactions On Yield.
Percent Yield

Quite often, the amount of product you actually get is less than the amount you calculated you should get; this is due to mechanical issues, incomplete reactions, “side” reactions, etc. How do you quantify the efficiency of a reaction?

\[
% \text{ Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%
\]

So, if the theoretical yield is 12.1 g and you actually isolate 9.92 g, your % yield is 82.0% (3 s.f.)
Solutions and Composition

A *solution* is composed of a *solute* and a *solvent*:

- Solutions are homogeneous mixtures;
- The solute is *usually* the material present in the lower amount; it can also be thought of as the material that will take part in the chemical reaction you are studying;
- The solvent is the material into which the solute is dissolved and is *usually* present in the larger amount;

There are FOUR main ways to express solution composition:

- Mass percent
- Mole fraction
- Molarity
- molality
Concentrated vs. Dilute Solutions

• Concentrated Solution: one that has a high level of solute present.
• Dilute Solution: one that has a low level of solute present.
• Saturated Solution: a solution that contains the maximum amount of solute possible.
Converting a concentrated solution to a dilute solution.
Mass Percent

- Often used to express composition of concentrated solutions;
- Simply:

\[
\text{mass\%} = \frac{\text{mass solute}}{\text{mass solution}} \times 100\%
\]

\[
\text{mass\%} = \frac{\text{mass solute}}{\text{mass solvent} + \text{mass solute}} \times 100\%
\]

- Since mass/mass ratio, no temperature dependence
Mole Fraction

- Used when the number of particles ratio is important;
- Symbolized using \( x_A \); Simply:

\[
x_{\text{solute}} = \frac{\text{moles solute}}{\text{moles solute} + \text{moles solvent}}
\]

\[
x_{\text{solute}} = \frac{n_{\text{solute}}}{n_{\text{total}}}
\]

- Since mole/mole ratio, no temperature dependence;
- Sum of all the \( x_i \)'s must equal 1!
Molarity

• Most common concentration unit used in chemistry;

• Units are moles per liter; symbolized using \( M \):

\[
M = \frac{\text{moles solute}}{\text{Liters of solution}} = \frac{mol}{L}
\]

• Since volume term exists, temperature dependent;

• Note: denominator is liters of \textit{solution}, not solvent!
Calculating the Molarity of a Solution

Glycine \((H_2NCH_2COOH)\) is the simplest amino acid. What is the molarity of an aqueous solution that contains 0.715 mol of glycine in 495 mL?

**SOLUTION:**

\[
\frac{0.715 \text{ mol glycine}}{495 \text{ mL soln}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.44 \text{ M glycine}
\]
Calculating Mass of Solute in a Given Volume of Solution

A buffered solution maintains acidity as a reaction occurs. In living cells, phosphate ions play a key buffering role, so biochemists often study reactions in such solutions. How many grams of solute are in 1.75 L of 0.460 M sodium monohydrogen phosphate?

Molarity is the number of moles of solute per liter of solution. Knowing the molarity and volume leaves us to find the number of moles and then the number of grams of solute. The formula for the solute is Na$_2$HPO$_4$.

**SOLUTION:**

\[
1.75 \text{ L} \times \frac{0.460 \text{ mol}}{1 \text{ L}} = 0.805 \text{ mol Na}_2\text{HPO}_4
\]

\[
0.805 \text{ mol Na}_2\text{HPO}_4 \times \frac{141.96 \text{ g Na}_2\text{HPO}_4}{1 \text{ mol Na}_2\text{HPO}_4} = 114 \text{ g Na}_2\text{HPO}_4
\]
Preparing a Dilute Solution from a Concentrated Solution

Isotonic saline is a 0.15 M aqueous solution of NaCl that simulates the total concentration of ions found in many cellular fluids. Its uses range from a cleaning rinse for contact lenses to a washing medium for red blood cells. How would you prepare 0.80 L of isotonic saline from a 6.0 M stock solution?

It is important to realize the number of moles of solute does not change during the dilution but the volume does. The new volume will be the sum of the two volumes which is the total final volume.

**SOLUTION:**

\[
0.80 \text{ L soln} \times \frac{0.15 \text{ mol NaCl}}{\text{L soln}} = 0.12 \text{ mol NaCl}
\]

mol NaCl same in both solutions

\[
0.12 \text{ mol NaCl} \times \frac{6 \text{ mol NaCl}}{\text{L soln conc}} = 0.020 \text{ L soln}
\]
Calculating Amounts of Reactants and Products for a Reaction in Solution

Specialized cells in the stomach release HCl to aid digestion. If they release too much, the excess can be neutralized with antacids. A common antacid contains magnesium hydroxide, Mg(OH)₂, which reacts with the acid to form water and magnesium chloride solution. As a government chemist testing commercial antacids, you use 0.10 M HCl to simulate the acid concentration in the stomach. How many liters of “stomach acid” react with a tablet containing 0.10 g of magnesium hydroxide?

**Strategy:** Write a balanced equation for the reaction; find the grams of Mg(OH)₂; determine the mol ratio of reactants and products; use mols to convert to molarity.
Calculating Amounts of Reactants and Products for a Reaction in Solution: Solution

\[ \text{Mg(OH)}_2(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + 2\text{H}_2\text{O}(l) \]

\[ 0.10 \text{ g Mg(OH)}_2 \times \frac{\text{mol Mg(OH)}_2}{58.33 \text{ g Mg(OH)}_2} = 1.7 \times 10^{-3} \text{ mol Mg(OH)}_2 \]

\[ 1.7 \times 10^{-3} \text{ mol Mg(OH)}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg(OH)}_2} = 3.4 \times 10^{-3} \text{ mol HCl} \]

\[ 3.4 \times 10^{-3} \text{ mol HCl} \times \frac{1 \text{ L}}{0.10 \text{ mol HCl}} = 3.4 \times 10^{-2} \text{ L} \]
Molality

- Infrequently used; seen mostly in colligative property calculations and physical chemistry;
- Units are moles per kg solvent; symbolized using $m$:

$$m = \frac{\text{moles solute}}{\text{kg solvent}} = \frac{mol}{kg}$$

- Since no volume term exists, temperature independent;
- Note: denominator is kg of solvent, not solution!
- For dilute ($< 1M$) aqueous solutions, $M \approx m$
Summary of amount-mass-number stoichiometric relationships.
End Chapter 3