OUTLINE

1. Significant Figures
2. Dimensional Analysis
3. Elements and Atoms
4. Naming Compounds
5. The Mole: Atomic Mass, Molar Mass, and Avogadro’s Number
6. Chemical Composition: Mass Percent, Empirical/Molecular Formulas
7. Balancing Chemical Equations
8. Stoichiometry and Mole-Mole Ratios
9. Limiting Reactants & Yields
Counting Significant Figures

• Nonzero numbers are always significant

What about zeroes?

• Leading zeroes = not significant
• Captive/trapped zeroes = significant
• Trailing zeroes (after decimal) = significant

0.0003040 meters

not significant significant
Calculations with Significant Figures

ROUND AT THE END

Multiplication/Division:
number with less significant figures

Addition/Subtraction:
the less precise number
Calculations with Significant Figures

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Addition/Subtraction: the less precise number

Examples:

\[
\frac{2.7991}{4.22} = 0.663294
\]

\[
7.23 + 70 + 3.7795 = 81.0795
\]

\[
\frac{4.771 + 2.3}{3.12} = 2.26634615
\]
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Addition/Subtraction:
- the less precise number

Examples:

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Calculations with Significant Figures

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**Examples:**

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Back to the Basics: Fractions

We can do these types of calculations already

\[
\frac{4}{9} \times \frac{1}{6} = \frac{4}{54} = \frac{2}{27}
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\[
\frac{4}{5} \times \frac{3}{4} \times \frac{5}{2} = \frac{3}{2}
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Extend Fractions to Multiplying Units

Unit conversion is just multiplying fractions!

\[
\frac{4}{3} \times \frac{3}{4} \times \frac{5}{2} = \frac{3}{2}
\]

\[
\frac{\text{centimeter}}{\text{seconds}} \times \frac{\text{meter}}{\text{centimeter}} \times \frac{\text{kilometer}}{\text{meter}} = \frac{\text{kilometer}}{\text{second}}
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The Atom

<table>
<thead>
<tr>
<th>PARTICLE</th>
<th>MASS</th>
<th>CHARGE</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron</td>
<td>$9.11 \times 10^{-31}$ kg</td>
<td>1–</td>
</tr>
<tr>
<td>Proton</td>
<td>$1.67 \times 10^{-27}$ kg</td>
<td>1+</td>
</tr>
<tr>
<td>Neutron</td>
<td>$1.67 \times 10^{-27}$ kg</td>
<td>0</td>
</tr>
</tbody>
</table>

The nucleus is very dense: A proton/neutron is ~2000 times heavier than an electron.
Atomic Symbols

- **Mass Number** (A): Protons + Neutrons
- **Atomic Number** (Z): # of Protons = # of Electrons
- **Number of Neutrons**: A - Z

Example: **Na**

- Mass Number: 23
- Atomic Number: 11
- Number of Neutrons: 23 - 11 = 12
- Protons: 11
- Electrons: 11

{\text{Na}_{11}^{23}}
**Atomic Symbols: Practice Problems**

<table>
<thead>
<tr>
<th>SYMBOL</th>
<th>$^{64}_{30}$Zn</th>
<th>$^{32}_{16}$S</th>
</tr>
</thead>
<tbody>
<tr>
<td># Protons</td>
<td>40</td>
<td></td>
</tr>
<tr>
<td># Neutrons</td>
<td>50</td>
<td></td>
</tr>
<tr>
<td># Electrons</td>
<td>40</td>
<td></td>
</tr>
<tr>
<td>Mass Number</td>
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**Mass Number (A)**

- Protons + Neutrons

**Atomic Number (Z)**

- # of protons
- # of electrons

**# Neutrons**

$\text{A} - \text{Z}$
# Atomic Symbols: Practice Problems

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**MASS NUMBER (A)**

protons + neutrons

**ATOMIC NUMBER (Z)**

# of protons

# of electrons

# NEUTRONS

= A – Z
# Atomic Symbols: Practice Problems

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**MASS NUMBER (A)**

protons + neutrons

**# NEUTRONS**

$= A - Z$

**ATOMIC NUMBER (Z)**

# of protons

# of electrons

charge
# Atomic Symbols: Practice Problems

## Elements and Atoms

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**MASS NUMBER (A)**
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\(A - Z\)

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## Atomic Symbols: Practice Problems

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**Mass Number (A)**: protons + neutrons

**Atomic Number (Z)**: # of protons

# of electrons

# Neutrons = A – Z

Charge

$\text{Symbol} \quad \begin{array}{c} \text{Mass Number (A)} \\ \text{protons + neutrons} \end{array}$
### General Rules for Naming Compounds

**IONIC**
- Metal + Nonmetal
- Cation + Anion
- Must be **neutral** overall!

**Naming:**
- Cation + Anion Root + “-ide”
  - NaCl → Sodium Chloride

**MOLECULAR**
- Nonmetal + Nonmetal

**Naming:**
- 1\(^{\text{st}}\) element: full name
- 2\(^{\text{nd}}\) element: root + “-ide”
- Use prefixes (Table 2.2)
  - BF\(_3\) → Boron Trifluoride
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**More exotic rules**
- Transition metals require charge
  
  *Hint: Find charge of anion first!*
- Cation + Charge + Anion Root + “-ide”
  
  FeCl\(_2\) → Iron (II) Chloride
  PbO\(_2\) → Lead (IV) Oxide
- Polyatomics are “one ion” (Table 2.3)
  
  AgCN → Silver (I) Cyanide

**More exotic rules**
- Don’t use “mono-” for first atom
  
  NO → Nitrogen Monoxide
- Drop “extra” vowels
  
  N\(_2\)O\(_5\) → Dinitrogen Pentoxide
- Oxoanions: -ate has more O’s than -ite
  
  NO\(_3^-\) → Nitrate
  NO\(_2^-\) → Nitrite
How should I think about atomic mass?

- Atoms are very small and very light
- Periodic table reports the measured *average atomic mass*
  - 1 hydrogen atom = 1.008 amu (atomic mass unit)
  - Makes it really impractical to work with in a laboratory/life

*Convince yourself of this:*
The mass of one atom (even one molecule) is horribly small. For instance: 1 H atom weighs $1.674 \times 10^{-24}$ grams

Recognize this is **not** a useful unit to use in life!
How should I think about atomic mass?

• Ask yourself: If an amu is useless, what units do we want to use for mass? [Answer: grams]

• So, we would like if: H: 1.008 amu means 1.008 g
  C: 12.01 amu means 12.01 g

• But how…?
The Mole: Atomic Mass, Molar Mass, and Avogadro’s Number

If you placed 1 hydrogen atom onto a scale, it wouldn’t register. If you placed 2 hydrogen atoms onto a scale, it still wouldn’t register. Now, imagine placing hydrogen atoms until you get 1.008 g.

If you placed 1 carbon atom onto a scale, it wouldn’t register. If you placed 2 carbon atoms onto a scale, it still wouldn’t register. Now, imagine placing carbon atoms until you get 12.01 g.
How should I think about atomic mass?

- Eventually you will put enough atoms to achieve a reading of 1.008 g for hydrogen and 12.01 g for carbon. In other words:

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• But how many atoms (“x”) did we need to get that mass in g?
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• But how many atoms (“x”) did we need to get that mass in g?

$$1.008 \text{ g } H \times \frac{1 \text{ atom } H}{1.674 \times 10^{-24} \text{ g}} = 6.022 \times 10^{23} \text{ atoms } H$$

Look familiar?!
Avogadro’s Number = \( 6.022 \times 10^{23} \) and the MOLE

• Avogadro’s number is what connects the atomic masses (in amu) to masses in grams! So…

\[
\begin{align*}
1.008 \text{ g H} &= 6.022 \times 10^{23} \text{ atoms H} = 1 \text{ mol H} \\
12.01 \text{ g C} &= 6.022 \times 10^{23} \text{ atoms C} = 1 \text{ mol C} \\
16.00 \text{ g O} &= 6.022 \times 10^{23} \text{ atoms O} = 1 \text{ mol O} \\
14.01 \text{ g N} &= 6.022 \times 10^{23} \text{ atoms N} = 1 \text{ mol N}
\end{align*}
\]

• Take-home message: The units on the periodic table that are useful are g/mol (molar mass). Often, we are dealing with large numbers of atoms/molecules, so g/mol is easier to deal with.
How many atoms of Mg are in 10.0 g Mg?
How many atoms of Mg are in 10.0 g Mg?

The molar mass of Mg is 24.31 g/mol.
Let's do it in steps just for clarity.

First, convert mass to moles using molar mass of Mg:

\[
10.0 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 0.411_4 \text{ mol Mg}
\]

Second, convert moles to atoms using Avogadro's number:

\[
0.411_4 \text{ mol Mg} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 2.48 \times 10^{23} \text{ atoms Mg}
\]

Everything in one giant conversion:

\[
10.0 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 2.48 \times 10^{23} \text{ atoms Mg}
\]
SUMMARY: MOLE IS CENTRAL

<table>
<thead>
<tr>
<th># ATOMS</th>
<th># MOLECULES</th>
<th># MOLES</th>
</tr>
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</table>

use: \( \frac{6.022 \times 10^{23}}{1 \text{ mol}} \) atoms or \( \frac{1 \text{ mol}}{6.022 \times 10^{23}} \) atoms

use molar mass from periodic table: \( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \) or \( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \)
Which has more atoms?

2 kg of Cu or 2 kg Mg
Which has more atoms?

2 kg of Cu or 2 kg Mg

The atomic mass of Cu is 63.55 amu.
The atomic mass of Mg is 24.31 amu.

→ Since each atom of Mg weighs less (1 atom Mg = 24.31 amu), it will take more atoms of Mg to reach 2 kg of Mg.
Avogadro’s number also applies to molecules/compounds

• Similar to how we don’t really encounter the mass of one atom, we rarely encounter the mass of one molecule in real life.

• Again, these masses would be very small numbers.

• For compounds or molecules, we also work in moles.

• The mass of 1 mole of a compound:

  Ex.) The molar mass of $\text{H}_2\text{O}$ is 18.02 g/mol.

\[
\begin{align*}
1 \text{ molecule } \text{H}_2\text{O} & \rightarrow 2 \text{ atoms } \text{H} + 1 \text{ atom } \text{O} \\
1 \text{ mol } \text{H}_2\text{O} & \rightarrow 2 \text{ mol H} + 1 \text{ mol O} \\
1 \text{ mol } \text{H}_2\text{O} & = 2 \cdot (1.008 \text{ g}) + 1 \cdot (16.00 \text{ g}) = 18.02 \text{ g } \text{H}_2\text{O}
\end{align*}
\]
Find the molar mass of ammonium carbonate.
Find the molar mass of ammonium carbonate.

First, write the formula: \((\text{NH}_4)_2\text{CO}_3\)

\[1 \text{ mol } (\text{NH}_4)_2\text{CO}_3 = 2 \text{ mol N} + 8 \text{ mol H} + 1 \text{ mol C} + 3 \text{ mol O}\]
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1 \text{ mol } (\text{NH}_4)_2\text{CO}_3 = 2 \text{ mol N} + 8 \text{ mol H} + 1 \text{ mol C} + 3 \text{ mol O}
\]

\[
\begin{align*}
2 \text{ mol N} & \rightarrow 2 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 28.02 \text{ g N} \\
8 \text{ mol H} & \rightarrow 8 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 8.064 \text{ g H} \\
1 \text{ mol C} & \rightarrow 1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C} \\
3 \text{ mol O} & \rightarrow 3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}
\end{align*}
\]

\[
96.09 \text{ g}
\]

\[
1 \text{ mol } (\text{NH}_4)_2\text{CO}_3 = 2 \text{ mol N} + 8 \text{ mol H} + 1 \text{ mol C} + 3 \text{ mol O} \\
= 2 \times (14.01 \text{ g}) + 8 \times (1.008 \text{ g}) + 1 \times (12.01 \text{ g}) + 3 \times (16.00 \text{ g}) \\
= 96.09 \text{ g}
\]
SUMMARY: MOLE IS CENTRAL

- Use: \( \frac{\text{6.022} \times 10^{23} \text{ molecules}}{1 \text{ mol}} \) or \( \frac{1 \text{ mol}}{\text{6.022} \times 10^{23} \text{ molecules}} \)

- Use molar mass of compound: \( \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \) or \( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \)
How many molecules are in a 50.0 g sample of ammonium carbonate?
How many molecules are in a 50.0 g sample of ammonium carbonate?

The molar mass of \((NH_4)_2CO_3\) is 96.09 g/mol.

The procedure is similar to what we did for atoms, but now use *molecules*.

First, convert from mass to number of moles using the molar mass:

\[
50.0 \text{ g} \ (NH_4)_2CO_3 \times \frac{1 \text{ mol} \ (NH_4)_2CO_3}{96.09 \text{ g} \ (NH_4)_2CO_3} = 0.5203 \text{ mol} \ (NH_4)_2CO_3
\]

Second, convert moles to number of *molecules* using Avogadro’s number:

\[
0.5203 \text{ mol} \ (NH_4)_2CO_3 \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 3.13 \times 10^{23} \text{ molecules} \ (NH_4)_2CO_3
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\text{mass} \times \frac{1 \text{ mol (NH}_4)_2\text{CO}_3}{96.09 \text{ g (NH}_4)_2\text{CO}_3} = \text{moles (NH}_4)_2\text{CO}_3
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\[
\text{moles (NH}_4)_2\text{CO}_3 \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \text{molecules (NH}_4)_2\text{CO}_3
\]

Everything in one giant conversion:

\[
50.0 \text{ g (NH}_4)_2\text{CO}_3 \times \frac{1 \text{ mol (NH}_4)_2\text{CO}_3}{96.09 \text{ g (NH}_4)_2\text{CO}_3} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 3.13 \times 10^{23} \text{ molecules (NH}_4)_2\text{CO}_3
\]
How many hydrogen atoms are in a 50.0 g sample of ammonium carbonate?

The molar mass of (NH₄)₂CO₃ is 96.09 g/mol.

The procedure is similar to what we did for atoms, but now use *molecules*.

\[
50.0 \text{ g } \text{(NH₄)}_2\text{CO₃} \times \frac{1 \text{ mol } \text{(NH₄)}_2\text{CO₃}}{96.09 \text{ g } \text{(NH₄)}_2\text{CO₃}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 3.13 \times 10^{23} \text{ molecules } \text{(NH₄)}_2\text{CO₃}
\]
How many hydrogen atoms are in a 50.0 g sample of ammonium carbonate?

The molar mass of \((\text{NH}_4)_2\text{CO}_3\) is 96.09 g/mol.
The procedure is similar to what we did for atoms, but now use *molecules*.

\[
50.0 \text{ g } (\text{NH}_4)_2\text{CO}_3 \times \frac{1 \text{ mol } (\text{NH}_4)_2\text{CO}_3}{96.09 \text{ g } (\text{NH}_4)_2\text{CO}_3} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 3.13 \times 10^{23} \text{ molecules } (\text{NH}_4)_2\text{CO}_3
\]

Remember that:
- 1 molecule of \((\text{NH}_4)_2\text{CO}_3\) = 2 atoms N + 8 atoms H + 1 atom C + 3 atoms O
- 1 mole \((\text{NH}_4)_2\text{CO}_3\) = 2 mole N + 8 mole H + 1 mole C + 3 mole O
How many hydrogen atoms are in a 50.0 g sample of ammonium carbonate?

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\]

Remember that:

• 1 molecule of \((\text{NH}_4)_2\text{CO}_3 = 2 \text{ atoms } \text{N} + 8 \text{ atoms } \text{H} + 1 \text{ atom } \text{C} + 3 \text{ atoms } \text{O}\)
• 1 mole \((\text{NH}_4)_2\text{CO}_3 = 2 \text{ mole } \text{N} + 8 \text{ mole } \text{H} + 1 \text{ mole } \text{C} + 3 \text{ mole } \text{O}\)

So:

\[
3.13 \times 10^{23} \text{ molecules } (\text{NH}_4)_2\text{CO}_3 \times \frac{8 \text{ atoms } \text{H}}{1 \text{ molecule } (\text{NH}_4)_2\text{CO}_3} = 2.51 \times 10^{24} \text{ atoms } \text{H}
\]
If you have equal mole samples of each compound, which contains the greatest number of oxygen atoms?

Magnesium nitrate

Dinitrogen pentoxide

Iron(III) phosphate

Barium oxide

Potassium acetate
If you have equal mole samples of each compound, which contains the greatest number of oxygen atoms?

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<td>Mg(NO₃)₂</td>
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<td>Mg(NO$_3$)$_2$</td>
<td>6 mol O</td>
</tr>
<tr>
<td>Dinitrogen pentoxide</td>
<td>N$_2$O$_5$</td>
<td>5 mol O</td>
</tr>
<tr>
<td>Iron(III) phosphate</td>
<td>FePO$_4$</td>
<td>4 mol O</td>
</tr>
<tr>
<td>Barium oxide</td>
<td>BaO</td>
<td>1 mol O</td>
</tr>
<tr>
<td>Potassium acetate</td>
<td>KCH$_3$CO$_2$</td>
<td>2 mol O</td>
</tr>
</tbody>
</table>
Introduction to mass percent

CHEMISTS CARE ABOUT MASS PERCENT!

\[
\% \text{ Mass} = \frac{\text{mass part}}{\text{mass whole}} \times 100\%
\]
How do I calculate the mass percentages for 1 mol CH₄?

• Remember that the molar mass of CH₄ is 16.04 g/mol:

\[ 1 \text{ mol CH}_4 = 1 \text{ mol C} + 4 \text{ mol H} \]
\[ = 1 (12.01 \text{ g}) + 4 (1.008 \text{ g}) = 16.04 \text{ g} \]

\[
% \text{ Mass} = \frac{\text{mass part}}{\text{mass whole}} \times 100\%
\]

% C → \( \frac{1(12.01) \text{ g}}{16.04 \text{ g}} \times 100\% = 74.90\% \text{ C} \)

% H → \( \frac{4(1.008) \text{ g}}{16.04 \text{ g}} \times 100\% = 25.10\% \text{ H} \)

100.0% total
How do I calculate the mass percentages for 2 mol CH₄?

• The molar mass of CH₄ is 16.04 g/mol, but now:

\[
2 \text{ mol CH}_4 = 2 \text{ mol C} + 8 \text{ mol H} \\
= 2(12.01 \text{ g}) + 8(1.008 \text{ g}) = 32.08 \text{ g}
\]

\[
\%	ext{ Mass} = \frac{\text{mass part}}{\text{mass whole}} \times 100\%
\]

\[
\%	ext{ C} \rightarrow \frac{2(12.01) \text{ g}}{32.08 \text{ g}} \times 100\% = 74.90\% \text{ C}
\]

\[
\%	ext{ H} \rightarrow \frac{8(1.008) \text{ g}}{32.08 \text{ g}} \times 100\% = 25.10\% \text{ H}
\]

\[
100.0\% \text{ total}
\]
Percent composition is **independent** of the starting amount!

This is why we usually assume we have 100 g or 1 mol. These are just super easy numbers to work with.

**Note:** If you wanted to use a strange amount, like 0.27 mol or 74.5 g of substance, your answers would be the same but the math isn’t as convenient. **BUT** you’ll still be right. 😊
What is the mass percent of nitrogen in barium nitrate?
What is the mass percent of nitrogen in barium nitrate?

First, write the formula: \( \text{Ba(NO}_3\text{)}_2 \)
Second, determine the molar mass = 261.35 g/mol
What is the mass percent of nitrogen in barium nitrate?

First, write the formula: \( \text{Ba(NO}_3\text{)}_2 \)

Second, determine the molar mass = 261.35 g/mol

We know that there are 2 mol N for 1 mol of \( \text{Ba(NO}_3\text{)}_2 \), so:

\[
\frac{2(14.01) \text{ g}}{261.35 \text{ g}} \times 100\% = 10.72\% \text{ N}
\]
If you have **equal** mass samples of each compound, which contains the **greatest** number of oxygen atoms?

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One way to solve this is to assume you have 100 g of each compound, then find the total number of O atoms in each compound.

This is a lot of work though!

Consider mass percentages!
If you have **equal** mass samples of each compound, which contains the **greatest** number of oxygen atoms?

<table>
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<tr>
<th>Compound</th>
<th>Formula</th>
<th>Mass Percent (%)</th>
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<tbody>
<tr>
<td>Magnesium nitrate</td>
<td>Mg(NO$_3$)$_2$</td>
<td>$\frac{6(16.00 \text{ g})}{148.33 \text{ g}} \times 100% = 64.72%$ O</td>
</tr>
<tr>
<td>Dinitrogen pentoxide</td>
<td>N$_2$O$_5$</td>
<td>$\frac{5(16.00 \text{ g})}{108.02 \text{ g}} \times 100% = 74.06%$ O</td>
</tr>
<tr>
<td>Iron(III) phosphate</td>
<td>FePO$_4$</td>
<td>$\frac{4(16.00 \text{ g})}{150.82 \text{ g}} \times 100% = 42.43%$ O</td>
</tr>
<tr>
<td>Barium oxide</td>
<td>BaO</td>
<td>$\frac{1(16.00 \text{ g})}{153.3 \text{ g}} \times 100% = 10.4%$ O</td>
</tr>
<tr>
<td>Potassium acetate</td>
<td>KCH$_3$CO$_2$</td>
<td>$\frac{2(16.00 \text{ g})}{98.14 \text{ g}} \times 100% = 32.61%$ O</td>
</tr>
</tbody>
</table>
Most often, we use mass percentages to help us figure out what compound we have.

These are called **EMPIRICAL FORMULAS**.
You have some “nitrogen oxide” compound and you want to figure out what it is (both formula and name). You know that it’s 46.7% nitrogen by mass.

• We want to know: \( N_xO_y = \) formula? name?

• Remember: the amount doesn’t matter for percent composition!
• Let’s assume we have 100 g of our \( N_xO_y \).
• This means that for every 100 g of \( N_xO_y \), we have:
  • 46.7 g of N
  • 53.3 g of O
You have some “nitrogen oxide” compound and you want to figure out what it is (both formula and name). You know that it’s 46.7% nitrogen by mass.

• We want to know: \( N_xO_y = \text{formula? name?} \)

We must convert the masses to \textit{moles}.

\[
\begin{align*}
N \rightarrow 46.7 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} &= 3.33 \text{ mol N} \\
O \rightarrow 53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} &= 3.33 \text{ mol O}
\end{align*}
\]
You have some “nitrogen oxide” compound and you want to figure out what it is (both formula and name). You know that it’s 46.7% nitrogen by mass.

- We want to know: \( \text{N}_x\text{O}_y = \text{formula? name?} \)

We need the simplest whole number ratio!

\[
\begin{align*}
\text{N} & \rightarrow 46.7 \, \text{g N} \times \frac{1 \, \text{mol N}}{14.01 \, \text{g N}} = 3.33 \, \text{mol N} \rightarrow 3.33 \, \text{mol N} \quad \text{\( \frac{3.33}{3.33} \) = 1 N} \\
\text{O} & \rightarrow 53.3 \, \text{g O} \times \frac{1 \, \text{mol O}}{16.00 \, \text{g O}} = 3.33 \, \text{mol O} \rightarrow 3.33 \, \text{mol O} \quad \text{\( \frac{3.33}{3.33} \) = 1 O}
\end{align*}
\]

Divide the number of moles by the SMALLEST value!
You have some “nitrogen oxide” compound and you want to figure out what it is (both formula and name).

You know that it’s 46.7% nitrogen by mass.

• We want to know: N_xO_y = formula? name?

We need the **simplest whole number ratio!**

\[
\begin{align*}
N \rightarrow & \quad 46.7 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 3.33 \text{ mol N} \rightarrow \frac{3.33 \text{ mol N}}{3.33} = 1 \text{ N} \\
O \rightarrow & \quad 53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.33 \text{ mol O} \rightarrow \frac{3.33 \text{ mol O}}{3.33} = 1 \text{ O}
\end{align*}
\]

Empirical Formula: NO | Nitrogen Monoxide
Aluminum oxide ($\text{Al}_x\text{O}_y$) is 41.51% Al and 36.92% O. Determine the empirical formula.

General procedure:

1. Assume a 100 g sample.

\[
\begin{align*}
\text{Al} & \rightarrow 41.51 \text{ g Al} \\
\text{O} & \rightarrow 36.92 \text{ g O}
\end{align*}
\]
Aluminum oxide ($\text{Al}_x\text{O}_y$) is 41.51% Al and 36.92% O. Determine the empirical formula.

General procedure:

1. Assume a 100 g sample.
2. Convert masses to moles.

\[
\begin{align*}
\text{Al} & \rightarrow 41.51 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 1.54 \text{ mol Al} \\
\text{O} & \rightarrow 36.92 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.31 \text{ mol O}
\end{align*}
\]
Aluminum oxide ($\text{Al}_x\text{O}_y$) is 41.51% Al and 36.92% O. Determine the empirical formula.

General procedure:

1. Assume a 100 g sample.
2. Convert masses to moles.
3. Divide the mole amounts by the smallest mole value.

$$\text{Al} \rightarrow 41.51 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 1.54 \text{ mol Al} \rightarrow \frac{1.54 \text{ mol Al}}{1.54} = 1 \text{ Al}$$

$$\text{O} \rightarrow 36.92 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.31 \text{ mol O} \rightarrow \frac{2.31 \text{ mol O}}{1.54} = 1.50 \text{ O}$$
Aluminum oxide \((\text{Al}_x\text{O}_y)\) is 41.51% Al and 36.92% O. Determine the empirical formula.

**General procedure:**

1. Assume a 100 g sample.
2. Convert masses to moles.
3. Divide the mole amounts by the smallest mole value.
4. Write empirical formula from simplest whole number ratio.

\[
\begin{align*}
\text{Al} & \quad \rightarrow 
\frac{41.51 \text{ g Al}}{26.98 \text{ g Al}} = 1.54 \text{ mol Al} \\
\text{O} & \quad \rightarrow 
\frac{36.92 \text{ g O}}{16.00 \text{ g O}} = 2.31 \text{ mol O}
\end{align*}
\]

\[
\begin{align*}
\text{Al} & \quad \rightarrow \quad \frac{1.54 \text{ mol Al}}{1.54} = 1 \text{ Al} \\
\text{O} & \quad \rightarrow \quad \frac{2.31 \text{ mol O}}{1.54} = 1.5 \text{ O}
\end{align*}
\]

Uh... these aren't whole numbers!
Aluminum oxide ($\text{Al}_x\text{O}_y$) is 41.51% Al and 36.92% O. Determine the empirical formula.

General procedure:
1. Assume a 100 g sample.
2. Convert masses to moles.
3. Divide the mole amounts by the smallest mole value.
4. Write empirical formula from simplest whole number ratio.

\[
\begin{align*}
\text{Al} & \quad \rightarrow \quad 41.51 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 1.54 \text{ mol Al} \quad \rightarrow \quad \frac{1.54 \text{ mol Al}}{1.54} = \quad 1 \text{ Al} \quad \times 2 \quad \rightarrow \quad 2 \text{ Al} \\
\text{O} & \quad \rightarrow \quad 36.92 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.31 \text{ mol O} \quad \rightarrow \quad \frac{2.31 \text{ mol O}}{1.54} = \quad 1.5 \text{ O} \quad \times 2 \quad \rightarrow \quad 3 \text{ O}
\end{align*}
\]

Multiply to get whole numbers!
Aluminum oxide (Al\textsubscript{x}O\textsubscript{y}) is 41.51% Al and 36.92% O. Determine the empirical formula.

**General procedure:**

1. Assume a 100 g sample.
2. Convert masses to moles.
3. Divide the mole amounts by the smallest mole value.
4. Write empirical formula from simplest whole number ratio.

\[
\begin{align*}
\text{Al} & \quad \rightarrow \quad 41.51 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 1.54 \text{ mol Al} \\
\text{O} & \quad \rightarrow \quad 36.92 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.31 \text{ mol O}
\end{align*}
\]

\[
\begin{align*}
\frac{1.54 \text{ mol Al}}{1.54} & \rightarrow 1 \text{ Al} \times 2 \rightarrow 2 \text{ Al} \\
\frac{2.31 \text{ mol O}}{1.54} & \rightarrow 1.50 \text{ mol O} \times 2 \rightarrow 3 \text{ O}
\end{align*}
\]

\[\text{Al}_2\text{O}_3\]
Calculate the empirical formula for a halohydrocarbon if it is
71.65% Cl, 24.27% C, and 4.07% H by mass.

Assuming a 100 g sample of the halohydrocarbon (Cl\textsubscript{a}C\textsubscript{b}H\textsubscript{c}):  

\[
\begin{align*}
\text{Cl} & \rightarrow 71.65 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 2.021 \text{ mol Cl} \rightarrow \frac{2.021 \text{ mol Cl}}{2.021} = 1 \text{ Cl} \\
\text{C} & \rightarrow 24.27 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.021 \text{ mol C} \rightarrow \frac{2.021 \text{ mol C}}{2.021} = 1 \text{ C} \\
\text{H} & \rightarrow 4.07 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.04 \text{ mol H} \rightarrow \frac{4.04 \text{ mol H}}{2.021} = 2 \text{ H}
\end{align*}
\]

The empirical formula is ClCH\textsubscript{2}. 
Determine the molecular formula for the same halohydrocarbon if it has a molar mass of 98.96 g/mol.

We determined previously that the empirical formula is ClCH₂. The empirical formula mass is 49.48 g/mol.

The molecular formula is always a multiple of the empirical formula. So: \( (\text{ClCH}_2)_n \)

We can determine the multiple \( (n) \) by taking the ratio between the molecular formula's molar mass and the empirical formula mass:

\[
 n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96 \text{ g}}{49.48 \text{ g}} = 2
\]

The molecular formula is \( (\text{ClCH}_2)_2 \) or \( \text{Cl}_2\text{C}_2\text{H}_4 \).
A compound containing only sulfur and nitrogen is 69.6% S by mass. If its molar mass is 184 g/mol, what is the correct name for it?

First, determine the empirical formula and empirical formula mass:

\[
\begin{align*}
    S & \rightarrow 69.6 \text{ g S} \times \frac{1 \text{ mol S}}{32.06 \text{ g S}} = 2.17 \text{ mol S} & \rightarrow & \frac{2.17 \text{ mol S}}{2.17} = 1 \text{ S} \\
    N & \rightarrow 30.4 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 2.17 \text{ mol N} & \rightarrow & \frac{2.17 \text{ mol N}}{2.17} = 1 \text{ N}
\end{align*}
\]

So the empirical formula is SN and the empirical mass is 46.07 g.

We can determine the multiple \((n)\) by taking the ratio between the molecular formula’s molar mass and the empirical formula mass:

\[
n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{184 \text{ g}}{46.07 \text{ g}} = 4
\]

The molecular formula is \((\text{SN})_4\) or \(\text{S}_4\text{N}_4\) or tetrasulfur tetranitride.
The peak farthest to the right (the one with the largest m/z value) tells you the estimated molecular mass of the compound.

Ex) Acetic Acid, C$_2$H$_4$O$_2$,
Molecular Mass = 60.05 g/mol

**Notes you don’t need to know:**
- The other peaks are broken fragments of the molecule.
- The height tells you the abundance of a fragment.
- We use mass spec. to identify elements with near-even (50%:50%) ratio of isotopes, like $^{79}$Br and $^{81}$Br.
What do chemical equations tell us?

- Formulas for the reactants (left side)
- Formulas for the products (right side)
- Phases, most of the time
- **Relative** amounts of reactants of reactants and products

\[ \text{REACTANT} \rightarrow \text{PRODUCTS} \]
What does it mean to be “balanced”?

- Same number of each type of atom on the left (reactants) and right (products) side.
- Law of Conservation of Mass

REACTANT \rightarrow \text{PRODUCTS}
How do we balance chemical equations?

• Mainly trial-and-error (some general strategies though).
• Make sure you have the same number of each type of atom on both sides of the equation.
• Do **NOT** balance by changing subscripts! Seriously, don’t.
• Balance the most complicated molecule *first*.

**REACTANT → PRODUCTS**
What makes a chemical reaction?

Hydrogen gas and oxygen gas react to form water vapor.

(words) hydrogen gas + oxygen gas → water vapor

(drawings) \( \infty \quad + \quad \text{O}_2 \quad \rightarrow \quad \text{H}_2 \text{O} \)

(equation) \( 2\text{H}_2 (g) \quad + \quad \text{O}_2 (g) \quad \rightarrow \quad 2\text{H}_2\text{O} (g) \)
How do I read a chemical equation?

- Subscripts are not conserved!
- Coefficients have no real meaning by themselves…
- RATIO of coefficient is what’s important.
- Read it like a recipe:
  “For every 2 H₂ molecules, we need 1 O₂ molecule to produce 2 H₂O molecules.”
Write a balanced chemical equation for ammonia synthesis from nitrogen and hydrogen gases.
Write a balanced chemical equation for ammonia synthesis from nitrogen and hydrogen gases.

Write out the core of the equation from the description:

\[ 1 \text{ N}_2 (g) + 3 \text{ H}_2 (g) \rightarrow 2 \text{ NH}_3 (g) \]

“To make 2 moles NH\textsubscript{3}, we need 1 mole N\textsubscript{2} and 3 moles H\textsubscript{2}.”
If 5.00 g of CH\textsubscript{4} (methane) is burned, what mass of water can be produced?
If 5.00 g of CH\(_4\) (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

\[ \underline{\text{____ CH}}_4 (g) \ + \underline{\text{____ O}}_2 (g) \ \rightarrow \underline{\text{____ CO}}_2 (g) \ + \underline{\text{____ H}}_2\text{O} (g) \]

**WHAT IS A COMBUSTION REACTION?**

When we “burn” a hydrocarbon (a compound with C, H, and/or O atoms), it *always* reacts with O\(_2\) gas in the air to form CO\(_2\) and H\(_2\)O gases as products.
If 5.00 g of CH$_4$ (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

\[ \text{1 CH}_4 (g) + \text{2 O}_2 (g) \rightarrow \text{1 CO}_2 (g) + \text{2 H}_2\text{O (g)} \]

“For every 1 mol CH$_4$, we need to react with 2 mol O$_2$ to produce 1 mol CO$_2$ and 2 mol H$_2$O.”
If 5.00 g of CH$_4$ (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

\[
1 \text{ CH}_4 (g) + 2 \text{ O}_2 (g) \rightarrow 1 \text{ CO}_2 (g) + 2 \text{ H}_2\text{O} (g)
\]

“For every 1 mol CH$_4$, we need to react with 2 mol O$_2$ to produce 1 mol CO$_2$ and 2 mol H$_2$O.”

**REMEMBER MY TIP:** If you don’t know how to start a problem, convert whatever they give you into **moles** first.
If 5.00 g of CH₄ (methane) is burned, what mass of water can be produced?

Write out the core of the equation from the description:

\[ \underline{1} \text{ CH}_4 (g) + \underline{2} \text{ O}_2 (g) \rightarrow \underline{1} \text{ CO}_2 (g) + \underline{2} \text{ H}_2\text{O} (g) \]

“For every 1 mol CH₄, we need to react with 2 mol O₂ to produce 1 mol CO₂ and 2 mol H₂O.”

**REMEMBER MY TIP:** If you don’t know how to start a problem, convert whatever they give you into **moles** first.

1. Use molar mass of CH₄ to convert from mass to moles.

\[
5.00 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} = 0.3117 \text{ mol CH}_4
\]
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2. Use 2:1 H₂O:CH₄ mole-mole ratio to find moles of H₂O.
   \[ 0.3117 \text{ mol CH}_4 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} = 0.6234 \text{ mol H}_2\text{O} \]
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   \]

3. Use molar mass of H₂O to convert from moles to mass.

   \[
   0.6234 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 11.2 \text{ g H}_2\text{O}
   \]
Pouring an aqueous solution of hydrochloric acid onto a solid block of magnesium metal produces an aqueous solution of magnesium chloride and hydrogen gas.

\[
\text{___} + \text{___} \rightarrow \text{___} + \text{___}
\]

A) Given 3.00 g Mg, how many moles of hydrochloric acid do we need?

B) If we produce 5.00 g H\textsubscript{2} gas, what mass of MgCl\textsubscript{2} solution is produced?

C) If we produce 4.00 g H\textsubscript{2} gas, what mass of HCl did we need?

\textbf{REMEMBER:} We only care about the ratio of coefficients, so we can use mole-mole ratios to go between reactants-to-reactants, reactants-products, products-to-reactants, or products-to-products.
Pouring an aqueous solution of hydrochloric acid onto a solid block of magnesium metal produces an aqueous solution of magnesium chloride and hydrogen gas.

\[
\begin{align*}
1 \text{ Mg (s)} & + 2 \text{ HCl (aq)} \rightarrow 1 \text{ MgCl}_2 \text{(aq)} & + & 1 \text{ H}_2 \text{(g)}
\end{align*}
\]

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A) Given 3.00 g Mg, how many moles of hydrochloric acid do we need?

\[
3.00 \text{ mol Mg} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg}} = 6.00 \text{ mol HCl}
\]

B) If we produce 5.00 g H\textsubscript{2} gas, what mass of MgCl\textsubscript{2} solution is produced?

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**REMEMBER:** We only care about the ratio of coefficients, so we can use mole-mole ratios to go between reactants-to-reactants, reactants-products, products-to-reactants, or products-to-products.
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3.00 \text{ mol Mg} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg}} = 6.00 \text{ mol HCl}
\]

B) If we produce 5.00 g H₂ gas, what mass of MgCl₂ solution is produced?

\[
5.00 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{1 \text{ mol MgCl}_2}{1 \text{ mol H}_2} \times \frac{95.21 \text{ g MgCl}_2}{1 \text{ mol MgCl}_2} = 236 \text{ g MgCl}_2
\]

C) If we produce 4.00 g H₂ gas, what mass of HCl did we need?

**REMEMBER:** We only care about the ratio of coefficients, so we can use mole-mole ratios to go between reactants-to-reactants, reactants-to-products, products-to-reactants, or products-to-products.
Pouring an aqueous solution of hydrochloric acid onto a solid block of magnesium metal produces an aqueous solution of magnesium chloride and hydrogen gas.

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

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C) If we produce 4.00 g H\textsubscript{2} gas, what mass of HCl did we need?

\[
4.00 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol H}_2} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 145 \text{ g HCl}
\]

**REMEMBER:** We only care about the ratio of coefficients, so we can use mole-mole ratios to go between reactants-to-reactants, reactants-to-products, products-to-reactants, or products-to-products.
I hope now you understand why I say to convert to moles before you do anything else. It’s because a balanced chemical equation gives us **mole-to-mole ratios** that we can use to convert between one reactant/product to another reactant/product.
2H₂ (g) + 1O₂ (g) → 2H₂O (g)

A) If you have 8 moles of hydrogen and all the oxygen you need, how many moles of water can you make?

\[ 8 \text{ mol H}_2 \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} = 8 \text{ mol H}_2\text{O} \]

B) If you have 6 moles of oxygen and all the hydrogen you need, how many moles of water can you make?

\[ 6 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} = 12 \text{ mol H}_2\text{O} \]

C) If you have 8 moles of hydrogen and 6 moles of oxygen, how many moles of water can you make?

Q: Which limits our reaction?

A: H₂ is limiting; 2 mol O₂ leftover

8 mol H₂O
When do I know if I need to figure out the limiting reactant?

When they give you the amounts of ALL reactants.
Limiting Reactants & Yield

METHOD 1

1. Assume one reactant is limiting and then determine amount of product you can form.
   \[ 8 \text{ mol } H_2 \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2} = 8 \text{ mol } H_2O \]

2. Assume other reactant is limiting and then determine amount of product you can form.
   \[ 6 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2} = 12 \text{ mol } H_2O \]

3. Reactant that limits amount of products formed is limiting reactant.
   \( H_2 \) produces less \( H_2O \) so it is limiting.

METHOD 2

1. Start with one reactant and determine how much of the other reactant you need.
   \[ 8 \text{ mol } H_2 \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2} = 4 \text{ mol } O_2 \]

2. Compare what you have vs. what you need:
   \[ \text{Have: } 6 \text{ mol } O_2 \]
   \[ \text{Need: } 4 \text{ mol } O_2 \]

3. We have more \( O_2 \) than we need
   \( \rightarrow O_2 \) excess
   \( \rightarrow H_2 \) is limiting.

CHOOSE THE METHOD THAT WORKS BEST FOR YOU!

\[ 2 \text{ H}_2 (g) + 1 \text{ O}_2 (g) \rightarrow 2 \text{ H}_2O (g) \]