Empirical vs. Molecular formula

- Last example shows difference between empirical and molecular formulas.
- **Empirical formula**: simplest possible formula with correct ratios of atoms
- **Molecular formula**: formula showing the actual composition of a molecule
- Can find molecular formula from empirical formula if we know molar mass

Formulas from % composition; more examples

- Hydrogen peroxide is 5.93% hydrogen and 94.07% oxygen by weight. What is its chemical formula?
- An unknown sample of a pure substance is 43.7% P and 56.3% O by weight. What is its chemical formula?

Note: mass ratios do not give mole ratios, since atomic masses are not the same.

Chemical Reactions

- Transformation of one or more chemical species into new substances.
- “reactants” → “products”
- Planning a synthesis is a “chemical design” problem.
Chemical Equations

• Written description of a reaction
• Varying levels of information: physical states, conditions, etc.
• Doesn’t necessarily mean that the reaction will take place readily, or at all.
  \[ \text{C (diamond)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \]

A Few Types of Reactions
(unbalanced equations!)

• “addition” or “combination” reactions
  \[ \text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6 \]
• “decomposition” reactions
  \[ \text{NH}_4\text{NO}_3 \rightarrow \text{N}_2 + \text{O}_2 + \text{H}_2\text{O} \]
• “substitution” or “displacement” reactions
  \[ \text{Mg} + \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \]

More Types of Reactions
(unbalanced equations!)

• “combustion” reactions (burning in \( \text{O}_2 \))
  \[ \text{C}_4\text{H}_{10} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]
• acid-base reactions
• precipitation reactions
• oxidation-reduction reactions

Balancing Equations: Conservation Laws

• In any chemical reaction, the following are conserved, and can be “accounted for.”
  ➔ number of atoms of each element
  ➔ mass
  ➔ energy
  ➔ electric charge

Meaning of “Balanced” Equations

2 \( \text{CO} + \text{O}_2 \rightarrow 2 \text{CO}_2 \)
• Coefficients give “reaction ratio”, and tell us how many ...
  ➔ Molecules react with molecules
  ➔ Moles react with moles
  ➔ NOT how many grams react with grams!!

Balancing Equations

• Find smallest whole number coefficients that satisfy conservation rules
• For many reactions, we do this “by inspection.” (trial and error)
• For some reactions, often use more systematic methods (redox reactions).
For example...

- Burning of propane \((C_3H_8)\) from a propane tank:

\[
C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O
\]

“Stoichiometry”

- Quantitative relationships in chemistry: “How much” or “How many” questions
- Applications of conservation laws
- “Composition stoichiometry” vs. “Reaction Stoichiometry”

Reaction Stoichiometry

- Balanced equation!
- “Moles react with moles.”
  - calculations centered on moles
- Use sample weight, molecular weight (molar mass), volume, density, etc. to relate known info to # of moles.

Example:

\[
2 C_2H_2 + 5 O_2 \rightarrow 4 CO_2 + 2 H_2O
\]

- How many grams of \(O_2\) are required to burn 52 g of \(C_2H_2\)?

Exhaust Problem

- First, formulate a plan of attack.
  - theory: How can we solve this?
  - Is it actually a chemistry problem?
  - data needed
  - assumptions to be made

“The average American car driven the average American distance in an average American year releases its own weight in carbon dioxide into the atmosphere.”

William McKittrick
The New Yorker
September 11, 1989

- Assess the current validity of this statement.
Data

- Assume gasoline = octane, C\textsubscript{8}H\textsubscript{18}
- Density of octane = 0.7 g/mL (or 0.7 kg/L)
- 1 gallon = 3785 mL
- 1 pound = 454 g (1 kg ~ 2.2 lbs.)
- Atomic weights: C = 12, O = 16, H = 1

Exhaust Problem

- Make decisions on remaining assumptions.
- Now use the data and your assumptions to reach a verdict. Is the statement reasonable?
- How confident are you in your answer?

US Car Fleet Mileage

Limiting Reagents

- In most reactions, the quantities mixed do not follow the stoichiometric ratio, so one reactant will run out before the others.
- “limiting reagent” vs. “in excess”
- In real applications, often choose limiting reagent based on cost considerations.

Limiting Reagents

2NO + O\textsubscript{2} → 2NO\textsubscript{2}

Which is the limiting reagent?
- O\textsubscript{2}
- NO
- NO\textsubscript{2}
- none

Limiting Reagents

2NO + O\textsubscript{2} → 2NO\textsubscript{2}
Example - Limiting Reagents

+ SiCl₄ is used in making computer chips. It is produced by the reaction:

SiO₂ + 2 C + 2 Cl₂ → SiCl₄ + 2 CO

+ How much SiCl₄ can be made from 75 g each of the reactants?

Reaction Yields

+ Most real reactions produce less product than equations would predict.
+ Competing reactions, impurities, time, etc.
+ Often report “percent yield”

Percent Yield

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

+ Actual yield: what you really get, from experiment
+ Theoretical yield: maximum you could expect, from calculations
+ actual yield < theoretical yield,
  \[\therefore% \text{ yield} < 100\%\]

Yield - example

+ Suppose we ran the reaction from the last example and obtained 80 g of SiCl₄.
What would our percent yield be?

Solutions & Concentration

Solutions

+ Homogeneous mixtures, 2 or more components
+ Solids, liquids, gases ...
+ **Solvent** = major component
+ **Solute** = minor component(s)
+ **Aqueous Solution** → water as solvent
Concentration
+ Solutions have “variable composition.”
+ To fully describe a solution, we need to know its concentration.
+ MOLARITY:
  # moles of solute / # liters of solution
+ Other concentration units also used

Example: Molarity
+ Cholesterol is C_{27}H_{46}O
+ “Cholesterol count” has units of mg per dL (1 dL = 10^{-1} L, or 100 mL)
+ Values above 200 are considered unhealthy.
  ➡️ Express this threshold in terms of molarity.

Ions & Ionic Compounds
+ Ions = atoms or groups of atoms with an electrical charge
+ Cations = positive charge
+ Anions = negative charge

Monatomic Ions
+ Formed from neutral atoms by adding or removing one or more electrons
+ Cations - metals, often from first 2 columns of the periodic table
  Na → Na^+ + e^-
+ Anions - nonmetals, often halogens
  Cl + e^- → Cl^-

Polyatomic Ions
+ 2 or more atoms
+ held together by chemical bonds
+ electrically charged
+ cations or anions
+ NH_4^+, H_3O^+, NO_3^-, SO_4^{2-}, etc.
  (Tables 2.5 & 2.6, p. 55, list several more.)
  LEARN ALL OF TABLE 2.5 & 2.6!

Ionic Compounds
+ Combine anions and cations, with charges balancing
+ Held together by Coulomb force (attraction of opposite charges)
+ Ionic compounds are crystalline solids, with high melting points.
Solubility Rules (See Table 3.1)

1. The nitrates, chlorates and acetates of all metals are soluble in water. Silver acetate is sparingly soluble.
2. All sodium, potassium and ammonium salts are soluble in water.
3. The chlorides, bromides and iodides of all metals except lead, silver and mercury(I) are soluble in water. HgI₂ is insoluble in water. PbCl₂, PbBr₂, and PbI₂ are soluble in hot water. The water-insoluble chlorides, bromides and iodides are also insoluble in dilute acids.

Solubility Rules, cont.

4. The sulfates of all metals except lead, mercury(I), barium and calcium are soluble in water. Silver sulfate is slightly soluble. The water-insoluble sulfates are also insoluble in dilute acids.
5. The carbonates, phosphates, borates, sulfites, chromates and arsenates of all metals except sodium, potassium and ammonium are insoluble in water but soluble in dilute acids. MgCrO₄ is soluble in water; MgSO₃ is slightly soluble in water.

Solubility Rules

6. The sulfides of all metals except barium, calcium, magnesium, sodium, potassium and ammonium are insoluble in water. BaS, CaS and MgS are sparingly soluble.
7. The hydroxides of sodium, potassium, and ammonium are very soluble in water. The hydroxides of calcium and barium are moderately soluble. The oxides and hydroxides of all other metals are insoluble.