# Chapter 3. Stoichiometry: Calculations with Chemical Formulas and Equations

## Common Student Misconceptions

- Students confuse the subscripts in a chemical formula with the coefficients in front of the formula in a balanced reaction equation.
- Students have difficulties grasping the meaning of a mole as a "collective"; a mole of a substance contains a fixed number ($6.022 \times 10^{23}$) of "building blocks" (atoms for most elements, molecules for molecular substances, formula units for ionic substances) in the same fashion as a dozen means 12 (eggs, people, items, etc.).
- Students often do not understand that mass of 1 mole of substance X can be significantly different from the mass of substance Y.
- Some students cannot distinguish between the number of moles actually manipulated in the laboratory versus the number of moles required by stoichiometry.
- Students do not appreciate that the coefficients in an empirical formula are not exact whole numbers because of experimental or round-off errors. In general, students have problems with the existence of experimental error.
- Students do not understand the difference between the amount of material present in the laboratory (or given in the problem) and the number of moles required by stoichiometry.
- Students do not understand that the reagent that gives the smallest amount of product is the limiting reactant.
- Students are often quite happy with a percent yield in excess of 100%.

## Teaching Tips

- Students who have good high school backgrounds find this chapter quite easy. Others find this chapter extremely difficult. Very few students have heard the term *stoichiometry* and can be intimidated by the language of chemistry.
- Balancing equations requires some trial and error. Algorithm-loving students find this uncomfortable. It helps to make some "errors" on purpose when teaching how to balance reaction equations and show some common "tricks" of how to recover from those errors.
- The concept of limiting reactants is one of the most difficult for beginning students. Students need a lot of numerical practice here. The use of analogies is often quite helpful.
- Atomic weights may be obtained with many significant figures. Advise students to use a sufficient number of significant figures such that the number of significant figures in the answer to any calculation is not limited by the number of significant figures in the formula weights they use.

## Lecture Outline

### 3.1 Chemical Equations[1,2,3,4,5]

- The quantitative nature of chemical formulas and reactions is called **stoichiometry**.
- Lavoisier observed that mass is conserved in a chemical reaction.
  - This observation is known as the law of conservation of mass.

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[1] "More Chemistry in a Soda Bottle: A Conservation of Mass Activity" from Further Readings
[2] "Antoine Lavoisier and The Conservation of Matter" from Further Readings
[3] "Chemical Wastes and the Law of Conservation of Matter" from Further Readings
[4] "Formation of Water" Movie from Instructor's Resource CD/DVD
[5] Figure 3.3 from Transparency Pack

- **Chemical equations** give a description of a chemical reaction.
- There are two parts to any equation:
  - **reactants** (written to the left of the arrow) and
  - **products** (written to the right of the arrow):

$$2H_2 + O_2 \rightarrow 2H_2O$$

- There are two sets of numbers in a chemical equation:
  - numbers in front of the chemical formulas (called stoichiometric *coefficients*) and
  - numbers in the formulas (they appear as subscripts).
- Stoichiometric coefficients give the *ratio* in which the reactants and products exist.
- The subscripts give the ratio in which the atoms are found in the molecule.
  - Example:
    - $H_2O$ means there are <u>two H atoms</u> for each one molecule of water.
    - $\mathbf{2}H_2O$ means that there are <u>two water molecules</u> present.
- Note: in $2H_2O$ there are *four* hydrogen atoms present (two for each water molecule).

## Balancing Equations[6,7,8,9,10,11,12,13,14,15,16,17]

- Matter cannot be lost in any chemical reaction.
  - Therefore, the products of a chemical reaction have to account for all the atoms present in the reactants–we must *balance* the chemical equation.
  - When balancing a chemical equation we adjust the stoichiometric coefficients in front of chemical formulas.
    - Subscripts in a formula are *never* changed when balancing an equation.
    - Example: the reaction of methane with oxygen:

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

    - Counting *atoms* in the reactants yields:
      - 1 C;
      - 4 H; and
      - 2 O.
    - In the products we see:
      - 1 C;
      - 2 H; and
      - 3 O.
    - It appears as though H has been lost and O has been created.
    - To balance the equation, we adjust the stoichiometric coefficients:

$$CH_4 + \mathbf{2}O_2 \rightarrow CO_2 + \mathbf{2}H_2O$$

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[6] "Reading a Chemical Equation" Activity from Instructor's Resource CD/DVD
[7] Figure 3.4 from Transparency Pack
[8] "Balancing Chemical Equations by Inspection" from Further Readings
[9] "The Fruit Basket Analogy" from Further Readings
[10] "A New Inspection Method for Balancing Redox Equations" from Further Readings
[11] "On Balancing Chemical Equations: Past and Present (A Critical Review and Annotated Bibliography)" from Further Readings
[12] "Reading A Balanced Chemical Equation" Activity from Instructor's Resource CD/DVD
[13] "How to Say How Much: Amounts and Stoichiometry" from Further Readings
[14] Figure 3.5 from Transparency Pack
[15] "Counting Atoms" Activity from Instructor's Resource CD/DVD
[16] "Balancing Equations" Activity from Instructor's Resource CD/DVD
[17] "Sodium and Potassium in Water" Movie from Instructor's Resource CD/DVD

### Indicating the States of Reactants and Products

- The physical state of each reactant and product may be added to the equation:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

- Reaction conditions occasionally appear above or below the reaction arrow (e.g., "Δ" is often used to indicate the addition of heat).

***FORWARD REFERENCES***

- Stoichiometric coefficients will be used to determine molar ratios (stoichiometric factors) in stoichiometric questions later in Chapter 3 as well as in Chapter 4 (section 4.6 on solution stoichiometry), Chapter 5 (stoichiometry of heat and Hess's Law), Chapter 10 (stoichiometry of gaseous reactions), Chapter 20 (section 20.9 on electrolysis).
- Stoichiometric coefficients will appear as powers to which concentrations and pressures are raised when writing equilibrium constant expressions of reversible reactions in Chapters 15-17, 19-20, and when writing rate law equations for elementary steps in Chapter 14.

## 3.2 Some Simple Patterns of Chemical Reactivity

### Combination and Decomposition Reactions[18,19,20,21,22,23]

- In **combination reactions** two or more substances react to form one product.
- Combination reactions have more reactants than products.
  - Consider the reaction:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

    - Since there are fewer products than reactants, the Mg has combined with $O_2$ to form MgO.
    - Note that the structure of the reactants has changed.
    - Mg consists of closely packed atoms and $O_2$ consists of dispersed molecules.
    - MgO consists of a lattice of $Mg^{2+}$ and $O^{2-}$ ions.
- In **decomposition reactions** one substance undergoes a reaction to produce two or more other substances.
- Decomposition reactions have more products than reactants.
  - Consider the reaction that occurs in an automobile air bag:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

    - Since there are more products than reactants, the sodium azide has decomposed into sodium metal and nitrogen gas.

### Combustion Reactions

- **Combustion reactions** are rapid reactions that produce a flame.
  - Most combustion reactions involve the reaction of $O_2(g)$ from air.
  - Example: combustion of a hydrocarbon (propane) to produce carbon dioxide and water.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

***FORWARD REFERENCES***

- Combustion reactions will be mentioned in Chapter 5 (as exothermic reactions involving fuels) and further discussed in Chapter 24 (as oxidation of organic compounds).
- Additional specific/important types of reactions will be introduced throughout the textbook.

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[18] "Lime" from Further Readings
[19] Figure 3.6 from Transparency Pack
[20] "Reactions with Oxygen" Movie from Instructor's Resource CD/DVD
[21] "Air Bags" Animation from Instructor's Resource CD/DVD
[22] "Nitrogen Triiodide" Movie from Instructor's Resource CD/DVD
[23] "Formation of Aluminum Bromide" Movie from Instructor's Resource CD/DVD

## 3.3 Formula Weights

### Formula and Molecular Weights[24]

- **Formula weight** (FW) is the sum of atomic weights for the atoms shown in the chemical formula.
  - Example: FW ($H_2SO_4$)
    - = 2AW(H) + AW(S) + 4AW(O)
    - = 2(1.0 amu) + 32.1 amu + 4(16.0 amu) = 98.1 amu.
- **Molecular weight** (MW) is the sum of the atomic weights of the atoms in a molecule as shown in the molecular formula.
  - Example: MW ($C_6H_{12}O_6$)
    - = 6(12.0 amu) + 12 (1.0 amu) + 6 (16.0 amu)
    - = 180.0 amu.
- Formula weight of the repeating unit (*formula unit)* is used for ionic substances.
  - Example: FW (NaCl)
    - = 23.0 amu + 35.5 amu
    - = 58.5 amu.

### Percentage Composition from Formulas[25,26]

- *Percentage composition* is obtained by dividing the mass contributed by each element (number of atoms times AW) by the formula weight of the compound and multiplying by 100.

$$\%\ \text{element} = \frac{(\text{number of atoms of that element})(\text{atomic weight of element})(100)}{(\text{formula weight of compound})}$$

***FORWARD REFERENCES***

- The ability to calculate molecular and formula weights will be an essential skill in finding molar masses throughout the textbook.

## 3.4 Avogadro's Number and The Mole[27,28,29,30,31,32,33,34,35,36,37,38,39,40,41,42]

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[24] "Gram Formula Weights and Fruit Salad" from Further Readings
[25] "Percentage Composition and Empirical Formula—A New View" from Further Readings
[26] "Molecular Weight and Weight Percent" Activity from Instructor's Resource CD/DVD
[27] "Mole, Mole per Liter, and Molar: A Primer on SI and Related Units for Chemistry Students" from Further Readings
[28] "Developing an Intuitive Approach to Moles" from Further Readings
[29] "The Mole, the Periodic Table, and Quantum Numbers: An Introductory Trio" from Further Readings
[30] "The Size of a Mole" from Further Readings
[31] "What's a Mole For?" from Further Readings
[32] "The Mole Concept: Developing an Instrument to Assess Conceptual Understanding" from Further Readings
[33] "A Mole of M&Ms" from Further Readings
[34] "How to Visualize Avogadro's Number" from Further Readings
[35] "Measuring Avogadro's Number on the Overhead Projector" from Live Demonstrations
[36] "Demonstrations for Nonscience Majors: Using Common Objects to Illustrate Abstract Concepts" from Live Demonstrations
[37] "Using Monetary Analogies to Teach Average Atomic Mass" from Further Readings
[38] "Pictorial Analogies IV: Relative Atomic Weights" from Further Readings
[39] "Relative Atomic Mass and the Mole: A Concrete Analogy to Help Students These Abstract Concepts" from Further Readings

- The **mole** (abbreviated "mol") is a convenient measure of chemical quantities.
- 1 mole of something = $6.0221421 \times 10^{23}$ of that thing.
  - This number is called **Avogadro's number**.
  - Thus, 1 mole of carbon atoms = $6.0221421 \times 10^{23}$ carbon atoms.
- Experimentally, 1 mole of $^{12}C$ has a mass of 12 g.

## Molar Mass[43,44,45]

- The mass in grams of 1 mole of substance is said to be the **molar mass** of that substance. Molar mass has units of g/mol (also written $g \cdot mol^{-1}$).
- The mass of 1 mole of $^{12}C$ = 12 g. Exactly.
- The molar mass of a molecule is the sum of the molar masses of the atoms:
  - Example: The molar mass of $N_2$ = 2 × (molar mass of N).
- Molar masses for elements are found on the periodic table.
- The formula weight (in amu) is numerically equal to the molar mass (in g/mol).

## Interconverting Masses and Moles

- Look at units:
  - Mass: g
  - Moles: mol
  - Molar mass: g/mol
- To convert between grams and moles, we use the molar mass.

## Interconverting Masses and Numbers of Particles[46]

- Units:
  - Number of particles: $6.022 \times 10^{23}$ $mol^{-1}$ (Avogadro's number).
  - Note: g/mol × mol = g (i.e. molar mass × moles = mass), and
  - mol × $mol^{-1}$ = a number (i.e. moles × Avogadro's number = molecules).
- To convert between moles and molecules we use Avogadro's number.

***FORWARD REFERENCES***

- It may be desirable to calculate molar masses with higher precision in later chapters.
- Avogadro number will be used to calculate the energy of 1 mole of photons in Chapter 6.
- Moles of electrons will be used in Chapter 7 (ionization energy or electron affinity), and in Chapter 20 to balance half-reactions and solve electrolysis problems.
- Bond dissociation energies (Chapter 8) and tabulated thermodynamic data (Appendix C) are mostly expressed per mole of bonds or substance.
- Moles will be used to calculate *molar* and *molal* concentrations (Chapters 4 and 11).
- Moles will be used in the Ideal Gas Law in Chapter 10.

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[40] "Moles, Pennies, and Nickels" from Further Readings
[41] "A Mole Mnemonic" from Further Readings
[42] "Analogies to Indicate the Size of Atoms and Molecules and the Magnitude of Avogadro's Number" from Further Readings
[43] Figure 3.8 from Transparency Pack
[44] "Demonstrations of the Enormity of Avogadro's Number" from Further Readings
[45] "For Mole Problems, Call Avogadro: 602-1023" from Further Readings
[46] Figure 3.12 from Transparency Pack

## 3.5 Empirical Formulas from Analyses[47,48,49,50,51,52,53,54,55,56,57,58,59,60]

- Recall that the empirical formula gives the *relative* number of atoms of each element in the molecule.
- Finding empirical formula from mass percent data:
  - We start with the mass percent of elements (i.e. empirical data) and calculate a formula.
  - Assume we start with 100 g of sample.
  - The mass percent then translates as the number of grams of each element in 100 g of sample.
  - From these masses, the number of moles can be calculated (using the atomic weights from the periodic table).
  - The lowest whole-number ratio of moles is the empirical formula.
- Finding the empirical mass percent of elements from the empirical formula.
  - If we have the empirical formula, we know how many moles of each element is present in one mole of same.
  - Then we use molar masses (or atomic weights) to convert to grams of each element.
  - We divide the number of grams of each element by the number of grams of 1 mole of sample to get the fraction of each element in 1 mole of sample.
  - Multiply each fraction by 100 to convert to a percent.

### Molecular Formulas from Empirical Formulas[61]

- The empirical formula (relative ratio of elements in the molecule) may not be the molecular formula (actual ratio of elements in the molecule).
- Example: ascorbic acid (vitamin C) has the empirical formula $C_3H_4O_3$.
  - The molecular formula is $C_6H_8O_6$.
  - To get the molecular formula from the empirical formula, we need to know the molecular weight, MW.
  - The ratio of molecular weight (MW) to formula weight (FW) of the empirical formula must be a whole number.

### Combustion Analysis[62]

- Empirical formulas are routinely determined by combustion analysis.
- A sample containing C, H, and O is combusted in excess oxygen to produce $CO_2$ and $H_2O$.
- The amount of $CO_2$ gives the amount of C originally present in the sample.

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[47] "Water" 3-D Model from Instructor's Resource CD/DVD
[48] "Hydrogen Peroxide" 3-D Model from Instructor's Resource CD/DVD
[49] "Sucrose" 3-D Model from Instructor's Resource CD/DVD
[50] "Oxygen" 3-D Model from Instructor's Resource CD/DVD
[51] "Methane" 3-D Model from Instructor's Resource CD/DVD
[52] "Carbon Dioxide" 3-D Model from Instructor's Resource CD/DVD
[53] "Glycine (molecular form)" 3-D Model from Instructor's Resource CD/DVD
[54] "Ethane" 3-D Model from Instructor's Resource CD/DVD
[55] "Methanol" 3-D Model from Instructor's Resource CD/DVD
[56] "A Known-to-Unknown Approach to Teach About Empirical and Molecular Formulas" from Further Readings
[57] "Making Assumptions Explicit: How the Law of Conservation of Matter Can Explain Empirical Formula Problems" from Further Readings
[58] "A Simple Rhyme for a Simple Formula" from Further Readings
[59] Figure 3.13 from Transparency Pack
[60] "Empirical Formula Determination: $C_8H_6O$" Activity from Instructor's Resource CD/DVD
[61] "Copper Sulfate: Blue to White" from Live Demonstrations
[62] "Combustion of Hydrocarbons: A Stoichiometry Demonstration" from Live Demonstrations

- The amount of $H_2O$ gives the amount of H originally present in the sample.
  - Watch the stoichiometry: 1 mol $H_2O$ contains 2 mol H.
- The amount of O originally present in the sample is given by the difference between the amount of sample and the amount of C and H accounted for.
- More complicated methods can be used to quantify the amounts of other elements present, but they rely on analogous methods.

## 3.6 Quantitative Information from Balanced Equations[63,64,65,66,67,68,69,70]

- The coefficients in a balanced chemical equation give the relative numbers of molecules (or formula units) involved in the reaction.
- The stoichiometric coefficients in the balanced equation may be interpreted as:
  - the relative numbers of molecules or formula units involved in the reaction or
  - the relative numbers of moles involved in the reaction.
- The molar quantities indicated by the coefficients in a balanced equation are called *stoichiometrically equivalent quantities*.
- Stoichiometric factors (or molar ratios) may be used to convert between quantities of reactants and products in a reaction.
- It is important to realize that the stoichiometric ratios are the ideal proportions in which reactants are needed to form products.
- A balanced reaction equation often provides more stoichiometric factors (or molar ratios) than needed to solve any particular stoichiometric problem. Often only one or two of them are relevant in a given problem.
- The number of grams of reactant cannot be *directly* related to the number of grams of product.
  - To get grams of product from grams of reactant:
    - convert grams of reactant to moles of reactant (use molar mass),
    - convert moles of one reactant to moles of other reactants and products (use the stoichiometric ratio from the balanced chemical equation),
    - convert moles back into grams for desired product (use molar mass).

***FORWARD REFERENCES***

- In Chapter 4 students will learn how to convert solution molarity and volume data into moles.
- In Chapter 10 students will learn how to use P, V and T information to find moles of gas.
- Stoichiometry of reactions will be further exploited when writing rate law expressions for elementary steps (Chapter 14) as well as equilibrium constant and reaction quotient expressions (Chapters 15, 16, 17, 19, and 20).
- Acid-base titrations mentioned in Chapter 4, and further discussed in Chapter 17, are practical applications of stoichiometry of acid-base neutralization reactions.

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[63] "How Many Digits Should We Use in Formula or Molar Mass Calculation" from Further Readings
[64] "Amounts Tables as a Diagnostic Tool for Flawed Stoichiometric Reasoning" from Further Readings
[65] "Stoogiometry: A Cognitive Approach to Teaching Stoichiometry" from Further Readings
[66] "Teaching Stoichiometry: A Two-Cycle Approach" from Further Readings
[67] "Pictorial Analogies XII: Stoichiometric Calculations" from Further Readings
[68] Figure 3.16 from Transparency Pack
[69] "Stoichiometry Calculation" Activity from Instructor's Resource CD/DVD
[70] "A Recipe for Teaching Stoichiometry" from Further Readings

## 3.7 Limiting Reactants[71,72,73,74,75,76,77,78,79,80,81]

- It is not necessary to have all reactants present in stoichiometric amounts.
- Often, one or more reactants is present in excess.
- Therefore, at the end of reaction those reactants present in excess will still be in the reaction mixture.
- The one or more reactants which are completely consumed are called the **limiting reactants**.
  - Reactants present in excess are called *excess reactants*.
- Consider 10 $H_2$ molecules mixed with 7 $O_2$ molecules to form water.
  - The balanced chemical equation tells us that the stoichiometric ratio of $H_2$ to $O_2$ is 2 to 1:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

  - This means that our 10 $H_2$ molecules require 5 $O_2$ molecules (2:1).
  - Since we have 7 $O_2$ molecules, our reaction is *limited* by the amount of $H_2$ we have (the $O_2$ is present in excess).
  - So, all 10 $H_2$ molecules can (and do) react with 5 of the $O_2$ molecules producing 10 $H_2O$ molecules.
  - At the end of the reaction, 2 $O_2$ molecules remain unreacted.

### Theoretical Yields

- The amount of product predicted from stoichiometry, taking into account limiting reactants, is called the **theoretical yield**.
  - This is often different from the *actual yield* – the amount of product actually obtained in the reaction.
- The **percent yield** relates the actual yield (amount of material recovered in the laboratory) to the theoretical yield:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

***FORWARD REFERENCES***

- Buffering action calculations in Chapter 17 (section 17.2) and pH calculations in acid-base titrations (section 17.3) can be viewed as stoichiometric problems with a limiting reactant (added strong acid or base).

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[71] "Learning Stoichiometry with Hamburger Sandwiches" from Further Readings
[72] "Limiting Reagent" Animation from Instructor's Resource CD/DVD
[73] "Limiting and Excess Reagents, Theoretical Yield" from Further Readings
[74] "Limiting Reagents" Activity from Instructor's Resource CD/DVD
[75] "Limiting Reactant: An Alternative Analogy" from Further Readings
[76] "Limiting Reagent Problems Made Simple for Students" from Further Readings
[77] "Election Results and Reaction Yields" from Further Readings
[78] Figure 3.17 from Transparency Pack
[79] "Interactive Demonstrations for the Mole Ratios and Limiting Reagents" from Live Demonstrations
[80] "Coffee, Coins, and Limiting Reagents" from Further Readings
[81] "A Dramatic Classroom Demonstration of Limiting Reagent Using the Vinegar and Sodium Hydrogen Carbonate Reaction" from Live Demonstrations

| ~~**3-D Models:**~~ | **Section:** |
|---|---|
| ~~Water~~ | 3.5 Empirical Formulas from Analyses |
| ~~Hydrogen Peroxide~~ | 3.5 Empirical Formulas from Analyses |
| ~~Sucrose~~ | 3.5 Empirical Formulas from Analyses |
| ~~Oxygen~~ | 3.5 Empirical Formulas from Analyses |
| ~~Methane~~ | 3.5 Empirical Formulas from Analyses |
| ~~Carbon Dioxide~~ | 3.5 Empirical Formulas from Analyses |
| ~~Glycine (molecular form)~~ | 3.5 Empirical Formulas from Analyses |
| ~~Methanol~~ | 3.5 Empirical Formulas from Analyses |
| ~~Ethane~~ | 3.5 Empirical Formulas from Analyses |

## Other Resources

| **Further Readings:** | **Section:** |
|---|---|
| More Chemistry in a Soda Bottle: A Conservation of Mass Activity | 3.1 Chemical Equations |
| Chemical Wastes and the Law of Conservation of Matter | 3.1 Chemical Equations |
| Antoine Lavoisier and the Conservation of Matter | 3.1 Chemical Equations |
| The Fruit Basket Analogy | 3.1 Chemical Equations |
| Balancing Chemical Equations by Inspection | 3.1 Chemical Equations |
| A New Inspection Method for Balancing Redox Equations | 3.1 Chemical Equations |
| On Balancing Chemical Equations: Past and Present (A Critical Review and Annotated Bibliography) | 3.1 Chemical Equations |
| How to Say How Much: Amounts and Stoichiometry | 3.1 Chemical Equations |
| Lime | 3.2 Some Simple Patterns of Chemical Reactivity |
| Gram Formula Weights and Fruit Salad | 3.3 Formula Weights |
| Percentage Composition and Empirical Formula–A New View | 3.3 Formula Weights |
| Mole, Mole per Liter, and Molar. A Primer on SI and Related Units for Chemistry Students | 3.3 Formula Weights |
| Using Monetary Analogies to Teach Average Atomic Mass | 3.4 Avogadro's Number and the Mole |
| Pictorial Analogies IV: Relative Atomic Weights | 3.4 Avogadro's Number and the Mole |
| Relative Atomic Mass and the Mole: A Concrete Analogy to Help Students These Abstract Concepts | 3.4 Avogadro's Number and the Mole |
| Moles, Pennies, and Nickels | 3.4 Avogadro's Number and the Mole |
| A Mole Mnemonic | 3.4 Avogadro's Number and the Mole |
| Analogies that Indicate the Size of Atoms and Molecules and the Magnitude of Avogadro's Number | 3.4 Avogadro's Number and the Mole |
| Developing an Intuitive Approach to Moles | 3.4 Avogadro's Number and the Mole |
| The Mole, the Periodic Table, and Quantum Numbers: An Introductory Trio | 3.4 Avogadro's Number and the Mole |
| The Size of a Mole | 3.4 Avogadro's Number and the Mole |
| What's a Mole For? | 3.4 Avogadro's Number and the Mole |
| The Mole Concept: Developing an Instrument | 3.4 Avogadro's Number and the Mole |

## Further Readings:

1. Frederic L. Holmes, "Antoine Lavoisier and The Conservation of Matter," *Chemical and Engineering News*, **September 12, 1994**, 38–45.

2. John W. Hill, "Chemical Wastes and the Law of Conservation of Matter," *J. Chem. Educ*., Vol. 58, **1981**, 996.

3. Daniel Q. Duffy, Stephanie A. Shaw, William D. Bare and Kenneth A. Goldsby, "More Chemistry in a Soda Bottle: A Conservation of Mass Activity," *J. Chem. Educ*., Vol. 72, **1995**, 734–736.

4. William C. Herndon, "On Balancing Chemical Equations: Past and Present (A Critical Review and Annotated Bibliography)," *J. Chem. Educ*., Vol. 74, **1997**, 1359–1362.

5. Zoltan Toth, "Balancing Chemical Equations by Inspection," *J. Chem. Educ*., Vol. 74, **1997**, 1363–1364.

6. William Bleam, Jr., "The Fruit Basket Analogy," *J. Chem. Educ*., Vol. 58, **1981**, 184. An analogy to help students master balancing equations.

7. Chunshi Guo, "A New Inspection Method for Balancing Redox Equations," *J. Chem. Educ*., Vol. 74, **1997**, 1365–1366.

8. Kenneth W. Watkins, "Lime," J. *Chem. Educ*., Vol. 60, **1983**, 60–63. An article on a some of the uses of quicklime (CaO) and hydrated lime ($Ca(OH)_2$).

9. Arthur M. Last and Michael J. Webb, "Using Monetary Analogies to Teach Average Atomic Mass," *J. Chem. Educ*., Vol. 70, **1993**, 234–235.

10. John H. Fortman, "Pictorial Analogies IV: Relative Atomic Weights," *J. Chem. Educ*., Vol. 70, **1993**, 235–236.

11. Josefina Arce de Sanabia, "Relative Atomic Mass and the Mole: A Concrete Analogy to Help Students Understand These Abstract Concepts," *J. Chem. Educ*., Vol. 70, **1993**, 233–234.

12. George L. Gilbert, "Percentage Composition and Empirical Formula–A New View," *J. Chem. Educ*., Vol. 75, **1998**, 851.

13. George Gorin, "Mole, Mole per Liter, and Molar. A Primer on SI and Related Units for Chemistry Students," *J. Chem. Educ*., Vol. 80, **2003**, 103–104.

14. Dawn M. Wakeley and Hans de Grys, "Developing an Intuitive Approach to Moles," *J. Chem. Educ.*, Vol. 77, **2000**, 1007–1009.

15. Mali Yin and Raymond S. Ochs, "The Mole, the Periodic Table, and Quantum Numbers: An Introductory Trio," *J. Chem. Educ.*, Vol. 78, **2001**, 1345–1347.

16. Miriam Toloudis, "The Size of a Mole," *J. Chem. Educ*., Vol. 73, **1996**, 348.

17. Sheryl Dominic, "What's a Mole For?" *J. Chem. Educ*., Vol. 73, **1996**, 309.

18. Shanthi R. Krishnan and Ann C. Howe, "The Mole Concept: Developing An Instrument to Assess Conceptual Understanding," *J. Chem. Educ*., Vol. 71, **1994**, 653–655.

19. R. Thomas Myers, "Moles, Pennies, and Nickels," *J. Chem. Educ*., Vol. 66, **1989**, 249.

20. Bernard S. Brown, "A Mole Mnemonic," *J. Chem. Educ*., Vol. 68, **1991**, 1039.

21. Carmela Merlo and Kathleen E. Turner, "A Mole of M&M's," *J. Chem. Educ*., Vol. 70, **1993**, 453.

22. Henk van Lubeck, "How to Visualize Avogadro's Number," *J. Chem. Educ*., Vol. 66, **1989**, 762.

23. Paul S. Poskozim, James W. Warrick, Permsook Tiempetpaisal and Joyce Albin Poskozim, "Analogies for Avogadro's Number," *J. Chem. Educ*., Vol. 63, **1986**, 125–126.

24. Damon Diemente, "Demonstrations of the Enormity of Avogadro's Number*," J. Chem. Educ*., Vol. 75, **1998**, 1565–1566.

25. R. E. Ulte, "For Mole Problems, Call Avogadro: 602-1023", *J. Chem. Educ.*, Vol. 79, **2002**, 1213.

26. M. Dale Alexander, Gordon J. Ewing, and Floyd T. Abbott, "Analogies that Indicate the Size of Atoms and Molecules and the Magnitude of Avogadro's Number," *J. Chem. Educ*., Vol. 61, **1984**, 591.

27. Wayne L. Felty, "Gram Formula Weights and Fruit Salad," *J. Chem. Educ*., Vol. 62, **1985**, 61.

28. P. K. Thamburaj, "A Known-to-Unknown Approach to Teach About Empirical and Molecular Formulas," *J. Chem. Educ.*, Vol. 78, **2001**, 915–916.

29. Stephen DeMeo, "Making Assumptions Explicit: How the Law of Conservation of Matter Can Explain Empirical Formula Problems," *J. Chem. Educ.*, Vol. 78, **2001**, 1050–1052.

30. Joel S. Thompson, "A Simple Rhyme for a Simple Formula," *J. Chem. Educ*., Vol. 65, **1988**, 704. An easy way to remember the strategy for converting percentage composition to an empirical formula. "Percent to mass, Mass to mol, Divide by small, Multiply 'til whole".

31. Christer Svensson, "How Many Digits Should we Use in Formula or Molar Mass Calculation", *J. Chem. Educ*., Vol. 81, **2004**, 827–829.

32. John Olmsted III, "Amounts Tables as a Diagnostic Tool for Flawed Stoichiometric Reasoning," *J. Chem. Educ.*, Vol. 76, **1999**, 52–54.

33. Carla R. Krieger, "Stoogiometry: A Cognitive Approach to Teaching Stoichiometry," *J. Chem. Educ*., Vol. 74, **1997**, 306–309.

34. Richard L. Poole, "Teaching Stoichiometry: A Two-Cycle Approach," *J. Chem. Educ*., Vol. 66, **1989**, 57.

35. Jean B. Umland, "A Recipe for Teaching Stoichiometry," *J. Chem. Educ*., Vol. 61, **1984**, 1036–1037.

36. Addison Ault, "How to Say How Much: Amounts and Stoichiometry," *J. Chem. Educ.*, Vol. 78, **2001**, 1347–1348.

37. John J. Fortman, "Pictorial Analogies XII: Stoichiometric Calculations," *J. Chem. Educ.*, Vol. 71, **1994**, 571–572.

38. Liliana Haim, Eduardo Corton, Santiago Kocmur, and Lydia Galagovsky, "Learning Stoichiometry with Hamburger Sandwiches," *J. Chem. Educ.*, Vol. 80, **2003**, 1021–1022.

39. Ernest F. Silversmith, "Limiting and Excess Reagents, Theoretical Yield," *J. Chem. Educ.*, Vol. 62, **1985**, 61.

40. Zoltan Toth, "Limiting Reactant: An Alternative Analogy," *J. Chem. Educ.*, Vol. 76, **1999**, 934.

41. A. H. Kalantar, "Limiting Reagent Problems Made Simple for Students," *J. Chem. Educ.*, Vol. 62, **1985**, 106.

42. Dennis McMinn, "Coffee, Coins, and Limiting Reagents," *J. Chem. Educ.*, Vol. 61, **1984**, 591.

43. Romeu C. Rocha-Filho, "Electron Results and Reaction Yields," *J. Chem. Educ.*, Vol. 64, **1987**, 248.

## Live Demonstrations:

1. Sally Solomon and Chinhyu Hur, "Measuring Avogadro's Number on the Overhead Projector," *J. Chem. Educ.*, Vol. 70, **1993**, 252–253. A monolayer of stearic acid on water is used to estimate Avogadro's number.

2. William Laurita, "Demonstrations for Nonscience Majors: Using Common Objects to Illustrate Abstract Concepts," *J. Chem. Educ.*, Vol. 67, **1990**, 60–61. This reference includes a demonstration of the measurement of Avogadro's number.

3. Lee. R. Summerlin,, Christie L. Borgford, and Julie B. Ealy, "Copper Sulfate: Blue to White," *Chemical Demonstrations, A Sourcebook for Teachers*, Volume 2 (Washington: American Chemical Society, **1988**), pp. 69–70. An exploration of color change associated with the dehydration of copper sulfate.

4. M. Dale Alexander and Wayne C. Woolsey, "Combustion of Hydrocarbons: A Stoichiometry Demonstration," *J. Chem. Educ.*, Vol. 70, **1993**, 327–328. The combustion of methane, propane, and butane are compared in this simple demonstration of stoichiometry.

5. Crystal Wood and Bryan Breyfogle, "Interactive Demonstrations for Mole Ratios and Limiting Reagents," *J. Chem. Educ.*, Vol. 83, **2006**, 741–748.

6. Romklao Artdej, Tienthong Thongpanchang, Stacy DeWees Hovede, "A Dramatic Classroom Demonstration of Limiting Reagent Using the Vinegar and Sodium Hydrogen Carbonate Reaction," *J. Chem. Educ.*, Vol. 85, **2008**, 1382-1384.