Chapter 3: Phenomena

Phenomena: When some substances are mixed together other substances form. Below is data for the reaction:

$$A(s) + 2B(aq) \rightarrow C(aq) + D(aq)$$

Look at the data below and identify any patterns that are present. Can you predict what would happen if you had 17.00 g A, 7.00 g B, 4.00 g C, and 5.00 g D?

<table>
<thead>
<tr>
<th>Initial</th>
<th>Final</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of A</td>
<td>0.00 g</td>
</tr>
<tr>
<td>Mass of B</td>
<td>20.00 g</td>
</tr>
<tr>
<td>Mass of C</td>
<td>0.00 g</td>
</tr>
<tr>
<td>Mass of D</td>
<td>20.00 g</td>
</tr>
<tr>
<td>Mass of D</td>
<td>0.00 g</td>
</tr>
<tr>
<td>Total</td>
<td>5.00 g</td>
</tr>
<tr>
<td>Total</td>
<td>13.97 g</td>
</tr>
<tr>
<td>Total</td>
<td>17.00 g</td>
</tr>
</tbody>
</table>

Significant Figures

Calculations with Significant Figures

- **Addition/Subtraction**: The answer is as precise as the least accurate number.
- **Multiplication/Division**: The number of significant figures in the answer is the same as number with the least amount of significant figures.

**Note**: If multiple operations are performed, follow the order of operations. Perform addition and subtraction first (with correct sig figs) then do multiplication and division (with correct sig figs).

<table>
<thead>
<tr>
<th>What Calculator would Report</th>
<th>What You Should Report (Correct Significant Figures)</th>
</tr>
</thead>
<tbody>
<tr>
<td>7.018</td>
<td></td>
</tr>
<tr>
<td>3.4223</td>
<td></td>
</tr>
<tr>
<td>0.7220769231</td>
<td>0.7221</td>
</tr>
</tbody>
</table>

Student Question

How many significant figures should appear in the answer for the following mathematical operation?

$$\frac{5.556 \times 2.30}{4.223 - 0.8}$$

a) 1  
b) 2  
c) 3  
d) 4  
e) None of the above

Significant Figures

- **Determining the Number of Significant Figures**
  - All non-zeros are significant
  - Interior zeros are significant
  - Leading zeros are not significant
  - Trailing zeros, after the decimal place, are significant

| Number of Significant Figures |  |
|------------------------------|  |
| 0.00004520090                 |  |
| 23,098,000                    |  |
| 200.                          |  |
Chapter 3: Stoichiometry

Molar Mass

- **Molar Mass (M):** The mass of a mole of objects.
  \[ M = \frac{m}{n} \]
  - where \( m \) is mass and \( n \) is moles.
- **Example:**
  - 1 mole of C weighs 12.01 grams
  - 12.01 g/mol
  - 12.01 amu

- **Average Atomic Mass:** The sum of the masses of an atom’s isotopes, each multiplied by its natural abundance.
  \[ \text{average atomic mass} = \frac{\text{fraction of isotope}_1 \times M_{\text{isotope}_1} + \text{fraction of isotope}_2 \times M_{\text{isotope}_2} + \ldots}{\text{fraction of isotope}_1 + \text{fraction of isotope}_2 + \ldots} \]

- **Note:** Plug in fraction of isotopes in decimals, not as a percentage.

Molar Mass of Compounds

- **Example:** Calculate the molar mass of \( \text{C}_2\text{H}_5\text{OH} \) (Ethanol):
  \[ M_{\text{C}_2\text{H}_5\text{OH}} = 2 \times M_{\text{C}} + 6 \times M_{\text{H}} + M_{\text{O}} \]
  \[ = 2 \times 12.01 \text{ amu} + 6 \times 1.01 \text{ amu} + 16.00 \text{ amu} \]
  \[ = 46.08 \text{ amu} \]

Conversions

- **Moles ↔ Atoms/Molecules/etc.:**
  - **Procedure:**
    1. Determine where you are starting and finishing
    2. Start calculation by writing down what you know
    3. Use dimensional analysis
      - Set up problem so you can cancel out units
  - **Tools Available:**
    - Avogadro’s Constant \( N_A = 6.022 \times 10^{23} \text{ mol}^{-1} \)
    - Periodic Table (M)
    - Chemical Formulas
    - Later Chemical Reactions

- **Example:**
  - How many moles in 5.23\( \times 10^{23} \) atoms?

- **Moles ↔ Grams:**
  - **Procedure:**
    1. Start calculation by writing down what you know
    2. Use the Periodic Table
  - **Example:**
    - How many moles of copper in 3.20 g?
      \[ M_{\text{Cu}} = 63.55 \text{ amu} \]

- **Moles ↔ Moles:**
  - **Procedure:**
    1. Start calculation by writing down what you know
    2. Use the Chemical Formula or Chemical Equation
  - **Example:**
    - How many moles of O in 3.4 mol of \( \text{CuSO}_4 \)?

Conversions Flow Chart

- **Conversion Flow Chart:**
  - **Tools:**
    - Grams
    - Molar Mass
    - Moles
    - Chemical Formulas
    - Later Chemical Reactions
  - **Procedure:**
    1. Determine where you are starting and finishing
    2. Start calculation by writing down what you know
    3. Use dimensional analysis
      - Set up problem so you can cancel out units
  - **Tools Available:**
    - Avogadro’s Constant \( N_A = 6.022 \times 10^{23} \text{ mol}^{-1} \)
    - Periodic Table (M)
    - Chemical Formulas
    - Later Chemical Reactions
Chapter 3: Stoichiometry

Chemical Equations

- **Chemical Reaction**: A process by which one or more substances are converted into other substances.
- **Reactants**: Starting material.
- **Products**: Substances formed.

Note: A chemical reaction is symbolized by an arrow.

- **Law of Conservation of Mass**: Mass is constant during a chemical reaction.

Note: The same number and type of atoms that are on the reactants side of the equation must be on the product side of the equation.

1. **Skeleton Equation (not balanced)**
   - \( \text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2 \)

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>NaOH</td>
</tr>
<tr>
<td>H</td>
<td>H_2</td>
</tr>
<tr>
<td>O</td>
<td></td>
</tr>
</tbody>
</table>

   - \( \text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow \text{NaOH(aq)} + \text{H}_2\text{(g)} \)

<table>
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<tbody>
<tr>
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</tr>
<tr>
<td>O</td>
<td></td>
</tr>
</tbody>
</table>

2. **Reaction Stoichiometry**: The quantitative relationship between the amounts of reactants consumed and those of products formed in chemical reactions as expressed by the balanced chemical equation for the reaction.

Example: \( \text{N}_2\text{(g)} + 3\text{H}_2\text{(g)} \rightarrow 2\text{NH}_3\text{(g)} \)

1 mole of \( \text{N}_2 \) and 3 moles of \( \text{H}_2 \) can make 2 moles of \( \text{NH}_3 \).

3. **Mole Ratios**: The stoichiometric relationship between two species in a chemical reaction written as a conversion factor.

Note: Equation must be balanced before constructing mole ratios.

4. **Writing Chemical Reactions**
   - **Step 1**: Write skeleton equation.
   - **Step 2**: Balance.
   - **Step 3**: Add state symbols
   - **Step 4**: Write other variables needed for the reaction to take place over the arrow.

Note: Never change subscripts to balance equations.

Examples: Solid (s), liquid (l), gas (g), aqueous (aq) compounds that are dissolved in water.

5. **Hints for Balancing Chemical Equations**
   - If an element occurs in only one compound on both sides of the equation, balance it first.
   - If an element occurs as a free element on either side of the chemical equation, balance it last.
   - When possible think of polyatomic ions as a group and balance them as 1 unit instead of individual atoms.

6. **Student Question**
   How many of the following statements are true concerning balanced chemical equations?
   i. The number of molecules is conserved.
   ii. Coefficients indicate mass ratios of the substances involved.
   iii. Atoms are neither created nor destroyed.
   iv. The sum of the coefficients on the left side equals the sum of the coefficients on the right side.

   a) 0  b) 1  c) 2  d) 3  e) 4

7. **Mole Ratios**: The stoichiometric relationship between two species in a chemical reaction written as a conversion factor.

Note: Equation must be balanced before constructing mole ratios.
**Limiting Reagents and Theoretical Yields**

2 bread + 1 bologna → 1 sandwich

What you have in your cupboard:
- 30 pieces of bread
- 7 pieces of bologna

<table>
<thead>
<tr>
<th>Initial</th>
<th>Bread</th>
<th>Bologna</th>
<th>Sandwich</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

---

**Limiting Reagents and Theoretical Yields**

**Student Question**

When Zn reacts with hydrochloric acid, zinc chloride and hydrogen gas form. If 25.0 g of zinc are react with 17.5 g of hydrochloric acid. How many grams of hydrogen gas will be produced?

Helpful Information:
- \( M_{\text{Zn}} = 65.41 \) g/mol
- \( M_{\text{HCl}} = 36.46 \) g/mol
- \( M_{\text{ZnCl}_2} = 100.86 \) g/mol
- \( M_{\text{H}_2} = 2.02 \) g/mol

a) 0.240 g  
   b) 0.385 g  
   c) 0.772 g  
   d) 0.970 g  
   e) None of the above

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**Limiting Reagents and Theoretical Yields**

- **Limiting Reagent**: The reactant that governs the maximum amount of product that can form.
- **Identifying The Limiting Reagent**
  - **Step 1**: Calculate the amount of each reactant in moles.
  - **Step 2**: Choose one of the reactants and use the mole ratio to calculate the theoretical amount of the second reactant needed. If \( n_{\text{cal}} < n_{\text{given}} \) then the reactant used for the calculation is the limiting reagent. If \( n_{\text{cal}} > n_{\text{given}} \) then the reactant not used for the calculation is the limiting reagent.
  - **Step 3**: Use an ICF table to determine the final number of moles of substances after the reaction has gone to completion.

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**Limiting Reagents and Theoretical Yields**

- **Theoretical Yield**: The maximum quantity (amount, mass, or volume) of product that can be obtained from a given quantity of a specified reactant.
- **Finding the Theoretical Yield**
  - **Step 1**: Identify the limiting reagent.
  - **Step 2**: Calculate theoretical yield based on the limiting reagent.

---

**Limiting Reagents and Theoretical Yields**

**Percent Yield**

- **Percent Yield**: The fraction of the theoretical yield actually produced.

\[
\text{Percent Yield} = \left( \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\%
\]

- If a student’s theoretical yield is 1.44 g and they only get 1.23 g. What is their percent yield?

---

**Limiting Reagents and Theoretical Yields**

**Determining Composition of Compounds**

- **Mass Percentage Composition**: The mass of each element expressed as a percentage of the total mass.

\[
\text{mass \% of element} = \left( \frac{m_{\text{element in the sample}}}{m_{\text{total}}} \right) \times 100\%
\]

**Note**: Mass percentage is an intrinsic property (independent of size).

- **Example**: A 3.16 g sample is made from 2.46 g of carbon, 0.373 g of hydrogen, and 0.329 g of oxygen. Determine the mass percentages of carbon.
Chapter 3: Stoichiometry

Determining Chemical Formulas

- **Empirical Formula:** Tells the relative number of atoms of each element present in the compound.
  
  Example: C\textsubscript{2}H\textsubscript{6}O\textsubscript{3} (glycerol) tells us that C, H, and O are present in a ratio of 2:6:3.

- **Molecular Formula:** Tells the actual numbers of atoms of each element in a molecule.
  
  Example: C\textsubscript{6}H\textsubscript{12}O\textsubscript{6} (glucose) tells us that exactly 6 C, 12 H, and 6 O atoms are in the compound.

Note: Different molecular formulas can have the same empirical formulas.

Example: C\textsubscript{2}H\textsubscript{4}O (formaldehyde) and C\textsubscript{2}H\textsubscript{4}O\textsubscript{2} (acetic acid) both have the empirical formula C\textsubscript{2}H\textsubscript{4}O.

Determining Chemical Formulas

**Determining Empirical Formulas**

- **Step 1:** Find number of grams of each atom. It is sometimes useful to assume that you have a 100 g sample.
- **Step 2:** Calculate moles of each type of atom (use molar mass).
- **Step 3:** Divide through by smallest number of moles calculated in step 2.
- **Step 4:** If necessary multiply all numbers in step 3 with a coefficient so that only whole numbers remain.

Take Away From Chapter 3

**Big Idea:** Depending on the situation it is convenient to talk about substances in terms of number of particles, mass, or moles. Given one of the quantities the other two can be calculated. A mole is just a specific number of objects similar to a dozen. Chemical equations allow us to find the relationship between reactants and products.

- **Significant Figures:**
  - Be able to determine the number of significant figures in a number.
  - Be able to perform operations (+,-,×, and ÷) and express answers with correct significant figures.

- **Moles:**
  - Know that a mole is just a specific number of objects (1 mol = 6.022×10\textsuperscript{23} objects) and can be used to classify anything.

- **Molar Mass:**
  - Know how to calculate the molar mass of an atom from relative abundances and atomic masses of its isotopes (23, 26, 27, 28).
  - Be able to calculate the molar mass of a compound.

- **Conversions:**
  - Be able to convert between the following:
    - Moles ↔ molecules/atoms (Avogadro's Number)
    - Grams ↔ moles (molar mass)
    - Moles ↔ moles (chemical formula or equation)

- **Chemical Equations:**
  - Be able to write and balance chemical equations.
  - Be able to determine the amount of products formed when given reactants.
Take Away From Chapter 3

- **Limiting Reagents and Theoretical Yields**
  - Be able to identify the limiting reagent (81, 82, 84, 137)
  - Be able to determine the theoretical yield
  - Be able to determine the amount of excess reactants.

- **Percent Yield**
  - \[ \text{Percent Yield} = \left( \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\% \]

- **Determining Composition of Compounds**
  - Be able to determine the mass percent of an element in a compound (41, 44, 47)
    - Usually easiest if assume sample size of 1 mol
  - \[ \text{mass} \% = \left( \frac{\text{mass of element}}{\text{mass of compound}} \right) \times 100\% \]

- **Determining Chemical Formulas**
  - Know the difference between empirical and molecular formulas.
    - Empirical: Relative number of atoms in a compound (usually for ionic compounds)
    - Molecular: The exact number of atoms in a molecule (molecular compounds)

- **Determining Chemical Formulas (Continued)**
  - Be able to determine empirical and molecular formulas from mass percentages of elements in compound (53, 54, 59)
  - Be able to determine empirical and molecular formulas when given combustion data (60, 61, 132)
    - \( \text{C}_x\text{H}_y\text{O}_z + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \) general combustion equation