INTRO AND BACKGROUND:
Reactions, Moles, Stoichiometry, and Solutions

Chemical Reaction

Atoms are REARRANGED to form a different substance

• Changes the way atoms are joined together
• Atoms CANNOT be created or destroyed!

Indicators of a Reaction

1) TEMPERATURE CHANGE
2) COLOR CHANGE
3) GAS FORMATION
4) PRECIPITATE

Symbols in Equations

• + separates the reactants or products
• → separates reactants from products
• ⇌ indicates a reversible reaction
• (s) solid
• (g) gas
• (l) liquid
• (aq) aqueous or water solution
• Δ indicates heat is supplied

CATALYST: speeds up reaction but is NOT a reactant or product

Balancing Chemical Equations

• Since we cannot break the Law of Conservation of Matter, equations MUST be balanced
• Balanced equations have the same number of each TYPE of atom on both sides of the equation
• COEFFICIENTS go in front of the formulas so the # of atoms of each element is the same on each side

WHAT GOES IN = WHAT COMES OUT!
Types of Reactions

- Millions of reactions exist... but there are only several categories of reactions.
- We will examine FIVE types:

1) **SYNTHESIS**
2) **DECOMPOSITION**
3) **SINGLE REPLACEMENT**
4) **DOUBLE REPLACEMENT**
5) **COMBUSTION**

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### Synthesis Reaction

*Two elements combine to make one compound*

**Combination Reaction**

\[ A + B \rightarrow AB \]

1) **BINARY COMPOUNDS** will break down into their elements

\[ 2 \text{NaCl} \rightarrow 2 \text{Na} + \text{Cl}_2 \]

### Decomposition Reaction

*One compound breaks down into two or more substances*

\[ AB \rightarrow A + B \]

2) **CARBONATES** break down into the oxide and \( \text{CO}_2 \)

\[ \text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2 \]
Decomposition Reaction

*One compound breaks down into two or more substances*

\[ AB \rightarrow A + B \]

3) **CHLORATES** break down into the binary salt and oxygen gas.

**EXAMPLE:**

\[ 2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2 \]

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Single-Replacement Reaction

*One element replaces another in a compound (element and compound become different element and compound)*

\[ A + BC \rightarrow B + AC \]

**EXAMPLE:**

Predict the products for the reaction and balance.

\[ \text{Zn} + 2 \text{HCl} \rightarrow \text{H}_2 + \text{ZnCl}_2 \]

**USE ACTIVITY SERIES TO SEE IF THEY HAPPEN!**

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Activity Series

**Activity Series of Metals**

<table>
<thead>
<tr>
<th>Metal</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>Li</td>
</tr>
<tr>
<td>Potassium</td>
<td>K</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg</td>
</tr>
<tr>
<td>Aluminum</td>
<td>Al</td>
</tr>
<tr>
<td>Zinc</td>
<td>Zn</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
</tr>
<tr>
<td>Mercury</td>
<td>Hg</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag</td>
</tr>
</tbody>
</table>

**Higher metal (more active) can REPLACE any metal lower than it, otherwise NO REACTION WILL OCCUR!!**

Metals from Li to Na will REPLACE H from acids and water... from Mg to Pb will REPLACE H from acids only!

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Double-Replacement Reaction

*Two elements replace one another*

\[ AB + CD \rightarrow AD + CB \]

**EXAMPLE:**

Predict the products for the reaction and balance.

\[ 3 \text{NaOH} + \text{FeCl}_3 \rightarrow 3 \text{NaCl} + \text{Fe(OH)}_3 \]

**USE SOLUBILITY RULES TO SEE IF THEY HAPPEN!**

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Combustion Reaction

*Compound with only C and H (sometimes O as well) is reacted with O₂*

\[ \text{C}_x\text{H}_y + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]

**EXAMPLE:**

Predict the products for the reaction and balance.

\[ 2 \text{C}_4\text{H}_{10} + 13 \text{O}_2 \rightarrow 8 \text{CO}_2 + 10 \text{H}_2\text{O} \]

**IF INCOMPLETE THEN CO AND H₂O FORM!**

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Conversion

*Mole conversions are useful but not practical in a lab...*

1 mole element = atomic mass (grams)

Get it right from the Periodic Table!!

**For example, 1 mole of arsenic has 74.9 g**
Molar Mass

Mass of 1 mole of an element, molecule, or compound

• How to Determine Molar Mass:
  1) Determine the # of moles of the individual elements that make up the compound (just look at the formula)
  2) Look up the mass of each element
  3) Multiply the mass of each by the # of moles of each
  4) Add up the masses

EXAMPLE:
Find the molar mass of glucose (C$_6$H$_{12}$O$_6$).

\[
\begin{align*}
6 \text{ C} & \times 12.00 \text{ g} = 72.00 \text{ g} \\
12 \text{ H} & \times 1.00 \text{ g} = 12.00 \text{ g} \\
6 \text{ O} & \times 16.00 \text{ g} = 96.00 \text{ g} \\
\hline
& = 180.00 \text{ g/mol}
\end{align*}
\]

Percent Composition

Amount or percent of an element in a compound by mass

• Determine the mass of each element and divide each by the total mass of the compound

\[
\text{Percent Composition} = \left( \frac{\text{Mass of element}}{\text{Mass of compound}} \right) \times 100
\]

Mole

Amount of something

• When measuring molecules and atoms, we use moles
• Used to count very small items
• Helps convert from the sub-microscopic to the macroscopic

BUT, WHAT AMOUNT?

6.02 x 10$^{23}$

"Avogadro’s Number"

That amount, but of what?

• REPRESENTATIVE PARTICLES: smallest piece of a chemical substance
  Ex: element = atom, ionic compound = formula unit, molecular compound = molecule

Conversions

1 mole = 6.02 x 10$^{23}$ atoms
1 mole = 6.02 x 10$^{23}$ molecules
1 mole = 6.02 x 10$^{23}$ formula units

These can be used in DIMENSIONAL ANALYSIS problems!!
Atoms to Moles

**EXAMPLE:**
A sample of Mg has $1.25 \times 10^{23}$ atoms of Mg. How many moles of Mg are contained in the sample?

\[
1.25 \times 10^{23} \text{ atoms Mg} \times \frac{1 \text{ mole of Mg}}{6.02 \times 10^{23} \text{ atoms Mg}} = \]

0.208 mol Mg

Practice

**EXAMPLES:**
- How many moles is 5.69 g NaOH?
- How many grams are there in 0.107 moles of CO₂?
- How many atoms are there in 12.21 g of C?

Mole-Volume Relationship

- Many chemicals exist as gases but difficult to mass
- Moles of a gas can be related to volume (Liters), but temperature and pressure also play a role
- Standard Temp. and Pressure (STP): 0°C and 1 atm
- At STP:
  \[1 \text{ mole gas} = 22.4 \text{ liters}\]

Empirical Formula

*Lowest whole number ratio of elements in a compound (can NOT be reduced)*

- **How to Determine:**
  1) Change the % to grams (if necessary)
  2) Convert grams to moles for each element
  3) Divide ALL of the mole answers by the SMALLEST # (mole ratio)
  4) If all whole #, then move on... if not then multiply to get whole # *SEE NEXT SLIDE!
  5) Use the whole # to represent the number of each element... write the formula

**EXAMPLE:**
Determine the empirical formula of the following compound: 12.12% C, 16.16% O, and 71.72% Cl.

\[
\begin{align*}
12.12 \text{ g C} \times \frac{1 \text{ mole C}}{12 \text{ g C}} &= 1.01 \times 1 = 1 \\
16.16 \text{ g O} \times \frac{1 \text{ mole O}}{16 \text{ g O}} &= 1.01 \times 1 = 1 \\
71.72 \text{ g Cl} \times \frac{1 \text{ mole Cl}}{35 \text{ g Cl}} &= 2.05 \times 1.01 = 2
\end{align*}
\]

**C₂H₂**
Ethane

**C₅H₈**
Styrene

**COCl₂**
Phosgene gas (used in WWI)

Empirical Formula

*Lowest whole number ratio of elements in a compound (can NOT be reduced)*

- After dividing ALL by the lowest #... If any are not whole numbers you must multiply ALL of the numbers by a number to make them all whole:

<table>
<thead>
<tr>
<th>IF YOU GET</th>
<th>MULTIPLY ALL BY</th>
</tr>
</thead>
<tbody>
<tr>
<td>.5</td>
<td>2</td>
</tr>
<tr>
<td>.3 or .6</td>
<td>3</td>
</tr>
<tr>
<td>.25</td>
<td>4</td>
</tr>
<tr>
<td>.20</td>
<td>5</td>
</tr>
</tbody>
</table>

Appeared on an old AP Exam!
Molecular Formula
Formula indicating the exact numbers of atoms of each element in the formula (can be reduced)

• How to Determine:
  1) Calculate the empirical formula (if needed)
  2) Calculate the molar mass of the empirical formula
  3) Divide the given molecular molar mass by the empirical molar mass
  4) Multiply subscripts of empirical formula by this #
  5) Write the molecular formula

EXAMPLE:
Determine the molecular formula of a compound composed of 85.7% C and 14.3% H with a molar mass of 70 g/mol.

\[
\begin{align*}
85.7 \text{ g C} & \times \frac{1 \text{ mole C}}{12 \text{ g C}} = 7.14 \\
14.3 \text{ g H} & \times \frac{1 \text{ mole H}}{1 \text{ g H}} = 14.3 \\
\end{align*}
\]

\[\text{CH}_2\]

Molecular Formula
Formula indicating the exact numbers of atoms of each element in the formula (can be reduced)

EXAMPLE:
Determine the molecular formula of a compound composed of 85.7% C and 14.3% H with a molar mass of 70 g/mol.

\[
\begin{align*}
1 \text{ C} & \times 12.00 \text{ g} = 12.00 \text{ g} \\
2 \text{ H} & \times 1.00 \text{ g} = 2.00 \text{ g} \\
\end{align*}
\]

\[\frac{12.00 \text{ g}}{14.00 \text{ g}} = 5 \]

\[\text{CH}_2 \rightarrow \text{C}_5\text{H}_{10}\]

Combustion Problems
Determine the amount of carbon, hydrogen, etc. present in the products formed which will tell you how much is present of each in the original sample (LAW OF CONSERVATION OF MASS) and then do the problems normally!

Practice
COMBUSTION EXAMPLE:
Combustion of 10.68 g of Vitamin C (containing only C, H, and O) yields 16.01 g of CO$_2$ and 4.37 g of H$_2$O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of this compound?

\[
\begin{align*}
16.01 \text{ g CO}_2 & \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} = 0.363 \text{ mol CO}_2 \\
4.37 \text{ g H}_2\text{O} & \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 0.242 \text{ mol H}_2\text{O} \\
12.01 \text{ g C} & \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1 \text{ mol C} \\
\end{align*}
\]

\[4.37 \text{ g C} + 0.489 \text{ g H} = 4.859 \text{ g} \]

\[10.68 \text{ g} - 4.859 \text{ g} = 5.821 \text{ g O}\]
**Practice**

**COMBUSTION EXAMPLE:**
Combustion of 10.68 g of Vitamin C (containing only C, H, and O) yields 16.01 g of CO$_2$ and 4.37 g of H$_2$O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of this compound?

Now you have the grams for each of the three elements in the compound... find the *empirical* and *molecular* as before!

\[
\begin{align*}
C_3H_4O_3 &= \text{Empirical} \\
C_6H_8O_6 &= \text{Molecular}
\end{align*}
\]

**Balanced Equations**

- Coefficients in a balanced chemical equation can represent a ratio of moles, molecules, liters (gases), or atoms... NOT GRAMS!
- Convert from an amount of one ingredient to another or to amounts of products
- Use MOLAR RATIO

Equations must be [CORRECTLY WRITTEN](#) and [BALANCED](#) in order to do these problems!!

**Stoichiometry Problems**

- Always follow this same basic format...

* * *

<table>
<thead>
<tr>
<th>Grams (Reactant)</th>
<th>Moles (Reactant)</th>
<th>Moles (Product)</th>
<th>Grams (Product)</th>
</tr>
</thead>
<tbody>
<tr>
<td><em>Use the MOLAR RATIO from the Equation</em></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><em>Problem can start or end with Moles and SKIP Grams altogether!</em></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[
\begin{array}{c}
\text{Grams R} \\
\times \quad 1 \text{ Mole R} \\
\times \quad \text{Moles P} \\
\times \quad \text{Grams P} \\
\text{Grams R} \\
\text{Moles R} \\
\text{1 Mole P}
\end{array}
\]

**Mass to Mass Conversions**

- **EXAMPLE:**
  If 10.1 g of Fe (3+) are added to a solution of copper (II) sulfate, how much solid copper would form?

\[
\begin{align*}
2 \text{ Fe} + 3 \text{ Cu}(\text{SO}_4) & \rightarrow 3 \text{ Cu} + \text{Fe}_2(\text{SO}_4)_3 \\
10.1 \text{ g Fe} & \times 1 \text{ mole Fe} \times 3 \text{ mole Cu} \times 64 \text{ g Cu} \\
& \times 56 \text{ g Fe} \times 2 \text{ mole Fe} \times 1 \text{ mole Cu}
\end{align*}
\]

\[17.3 \text{ g Cu}\]

**Mass to Volume Conversions**

- **EXAMPLE:**
  Potassium metal reacts with water to produce potassium hydroxide and hydrogen gas. If 12.1 g K is reacted completely, how many liters of H$_2$ gas can be produced at STP?

\[
\begin{align*}
2 \text{ K} + 2 \text{ H}_2\text{O} & \rightarrow 2 \text{ KOH} + \text{H}_2 \\
12.1 \text{ g K} & \times 1 \text{ mole K} \times 1 \text{ mole H}_2 \times 22.4 \text{ L H}_2 \\
& \times 39 \text{ g K} \times 2 \text{ mole K} \times 1 \text{ mole H}_2
\end{align*}
\]

\[3.47 \text{ L H}_2\]
Limiting Reagent

Reactant that runs out first in a chemical reaction

• Amount of reactants available for a reaction LIMITS the amount of product that can be made

• EXCESS REAGENT: reactant that is not used up completely

Limiting Reagent Problems

• EXAMPLE:
Copper reacts with sulfur to form copper (I) sulfide. If 10.6 g of Cu reacts with 3.83 g S, how much product will be formed?

\[ 2 \text{Cu} + \text{S} \rightarrow \text{Cu}_2\text{S} \]

\[
\begin{array}{c}
10.6 \text{ g Cu} \\
X \\
1 \text{ mole Cu} \\
X \\
64 \text{ g Cu} \\
X \\
2 \text{ mole Cu} \\
X \\
1 \text{ mole Cu}_2\text{S} \\
160 \text{ g Cu}_2\text{S}
\end{array} = 13.2 \text{ g} \\
\begin{array}{c}
3.83 \text{ g S} \\
X \\
1 \text{ mole S} \\
X \\
32 \text{ g S} \\
X \\
1 \text{ mole S} \\
X \\
1 \text{ mole Cu}_2\text{S} \\
19.2 \text{ g}
\end{array}
\]

Limiting Reagent Problems

• EXAMPLE:
How much of the excess reagent will be left over from the previous problem?

\[ 2 \text{Cu} + \text{S} \rightarrow \text{Cu}_2\text{S} \]

Sulfur is in EXCESS!

\[
\begin{array}{c}
10.6 \text{ g Cu} \\
X \\
1 \text{ mole Cu} \\
X \\
64 \text{ g Cu} \\
X \\
2 \text{ mole Cu} \\
X \\
1 \text{ mole S} \\
32 \text{ g S} \\
X \\
1 \text{ mole S} \\
X \\
1 \text{ mole Cu}_2\text{S} \\
19.2 \text{ g}
\end{array} = 2.65 \text{ g} \text{ needed} \text{ and } 3.83 \text{ g} - 2.65 \text{ g} = 1.18 \text{ g} \text{ excess}
\]

Percent Yield

Amount of product formed compared to the amount that should have formed

\[
\frac{\text{ACTUAL}}{\text{THEORETICAL}} \times 100
\]

• How to Determine:
  1) Actual is given or found in lab
  2) Calculate theoretical by dimensional analysis (may need limiting reagent)
  3) Use the equation

Should NEVER be greater than 100%... WHY?

Multiple Choice Practice

Reminder, there are NO CALCULATORS allowed on the Multiple Choice portion of the AP Chemistry Exam!!

• Need to do basic math problems (add / subtract / multiply / divide) by hand
• Use rounding / simplifying to make some problems easier
• Use logic to eliminate answers that do not make sense based on the numbers given

Multiple Choice Practice

• EXAMPLE:
When 8.0 g of N₂H₄ (32 g mol⁻¹) and 92 g of N₂O₄ (92 g mol⁻¹) are mixed together and react according to the equation below, what is the maximum mass of H₂O that can be produced?

\[ 2 \text{N}_2\text{H}_4(g) + \text{N}_2\text{O}_4(g) \rightarrow 3 \text{N}_2(g) + 4 \text{H}_2\text{O}(g) \]

(A) 9.0 g
(B) 18 g
(C) 36 g
(D) 72 g
(E) 144 g
Multiple Choice Practice

• **EXAMPLE:**
  When hafnium metal is heated in an atmosphere of chlorine gas, the product of the reaction is found to contain 62.2% Hf by mass and 37.4% Cl by mass. What is the **empirical formula** for this compound?
  
  (A) HfCl
  (B) HfCl₂
  (C) HfCl₃
  (D) HfCl₄
  (E) Hf₂Cl₃

Aqueous Solution

*Solution in which H₂O is the solvent*

• **Ionic and polar molecules dissolved best... WHY?**
• **MISCIBLE**: liquids completely mix (alcohol and water)
• **IMMISCIBLE**: don’t mix (oil and water)

Net Ionic Equations

*Shows only the ions directly involved in a chemical reaction*

• **SPECTATOR IONS**: appear in identical forms on both sides of the equation but are not directly involved in the reaction (cancel out)

• **Steps for Writing Net Ionic Equations:**
  1) Write the balanced equation
  2) Rewrite the equation showing the IONS that form in solution for each **SOLUBLE** compound
  3) Cancel spectator ions (ALL cancel = NO rxn!)
  4) Rewrite the final equation

Net Ionic Equations

*Shows only the ions directly involved in a chemical reaction*

• **EXAMPLE:**
  Pb(NO₃)₂ (aq) + 2 KI (aq) → PbI₂ (s) + 2 KNO₃ (aq)

Molarity (M)

*Moles of solute per liter of solution*

\[
M = \frac{\text{Moles of solute}}{\text{Liters of solution}}
\]

• **EXAMPLE:**
  Calculate the molarity of a solution when 11.5 g of NaOH is dissolved in enough water to make 1.5 liters of solution.

Preparing Solutions

• To make a certain concentration of solution, the **SOLID** should be weighed out first and then placed in a volumetric flask
• Dissolve the solid in **SOME** of the solvent then add the remaining solvent
**Example:**
What volume of 16 M H₂SO₄ must be used to prepare 1.5 L of 0.10 M solution?

\[(16 \text{ M})(V_1) = (0.10 \text{ M})(1.5 \text{ L})\]

\[0.0094 \text{ L} \quad \text{or} \quad 9.4 \text{ mL}\]

Add 1490.6 mL of water to make 0.10 M!

**Practice:**
How many mL of water should be used to prepare 300.0 mL of 0.750 M NaBr solution using 2.00 M stock solution?