Comprehensive Study Guide for Chemistry Exam 2
Chapters 5, 8, & 9

Chapter 5: Ionic and Covalent Compounds

1. Classifying Ionic vs. Molecular Compounds

- **Ionic Compounds** are formed between a **metal** (cation) and a **nonmetal** (anion). Electrons are transferred from the metal to the nonmetal, creating charged ions that are held together by electrostatic attraction. For example, NaCl is formed from the metal Na and the nonmetal Cl.
- **Molecular (Covalent) Compounds** are formed between two or more **nonmetals**. Electrons are shared between atoms to form covalent bonds. For example, H_2O is formed from the nonmetals H and O.

Practice Problems: Classify each of the following as ionic or molecular.

1. **Ionic Examples:**
   (a) $FeCl_3$: Iron (Fe) is a metal, Chlorine (Cl) is a nonmetal. $\rightarrow$ **Ionic**
   (b) K_2O: Potassium (K) is a metal, Oxygen (O) is a nonmetal. $\rightarrow$ **Ionic**
   (c) $Mg_3(PO_4)_2$: Magnesium (Mg) is a metal, and phosphate (${PO_4}^{3-}$) is a polyatomic anion (composed of nonmetals). The bond between Mg and the phosphate group is ionic. $\rightarrow$ **Ionic**
   (d) $Ca(OH)_2$: Calcium (Ca) is a metal, and hydroxide (OH^-) is a polyatomic anion. The bond between Ca and the hydroxide group is ionic. $\rightarrow$ **Ionic**
   (e) AlP: Aluminum (Al) is a metal, Phosphorus (P) is a nonmetal. $\rightarrow$ **Ionic**
2. **Molecular Examples:**
   (a) SO_2: Sulfur (S) and Oxygen (O) are both nonmetals. $\rightarrow$ **Molecular**
   (b) N_2O_4: Nitrogen (N) and Oxygen (O) are both nonmetals. $\rightarrow$ **Molecular**
   (c) PCl_5: Phosphorus (P) and Chlorine (Cl) are both nonmetals. $\rightarrow$ **Molecular**
   (d) CH_4: Carbon (C) and Hydrogen (H) are both nonmetals. $\rightarrow$ **Molecular**
   (e) $SiCl_4$: Silicon (Si) is a metalloid that bonds covalently with nonmetals like Chlorine (Cl). $\rightarrow$ **Molecular**

2. Lattice Energy

- **Lattice Energy** is the energy required to completely separate one mole of a solid ionic compound into its gaseous ions. It is a measure of the strength of the ionic bonds.
- **Trend Factors:**
  1. **Ionic Charge:** Lattice energy increases significantly as the magnitude of the ionic charges increases. For example, MgO ($+2, -2$) has a much higher lattice energy than NaCl ($+1, -1$).

2. **Atomic Radius (Ionic Size):** Lattice energy decreases as the size of the ions increases. Smaller ions can get closer together, resulting in stronger electrostatic attraction. For example, LiF (smaller ions) has a higher lattice energy than CsI (larger ions).

Practice Problems: Arrange the following ionic compounds in order of increasing lattice energy.

1. **NaCl, MgO, NaF**
   - Charges: NaCl $(+1, -1)$, NaF $(+1, -1)$, MgO $(+2, -2)$.
   - MgO has the highest charges, so it will have the highest lattice energy.
   - Both NaCl and NaF have the same charges. We compare ionic size. F^- is smaller than Cl^-. Therefore, NaF will have a higher lattice energy than NaCl.
   - **Order:** NaCl ¡ NaF ¡ MgO
2. **LiBr, KCl, MgS**
   - Charges: LiBr $(+1, -1)$, KCl $(+1, -1)$, MgS $(+2, -2)$.
   - MgS has the highest charges, so it will have the highest lattice energy.
   - LiBr and KCl have the same charges. Li^+ is smaller than K^+, and Br^- is larger than Cl^-. Comparing pairs, Li^+ and Br^- are closer in size to each other than K^+ and Cl^- are. Also, going down the periodic table decreases lattice energy. LiBr is higher on the table than KCl.
   - **Order:** KCl ¡ LiBr ¡ MgS
3. **AlN, MgS, LiF**
   - Charges: AlN $(+3, -3)$, MgS $(+2, -2)$, LiF $(+1, -1)$.
   - Lattice energy increases with the magnitude of the charges.
   - **Order:** LiF ¡ MgS ¡ AlN
4. **SrO, CaO, BaO**
   - Charges: All are $(+2, -2)$.
   - We must compare ionic radii. All have the same anion (O^{2-}). The cation size increases down the group: Ca ¡ Sr ¡ Ba.
   - Since lattice energy is inversely proportional to ionic size, the smallest cation will result in the highest lattice energy.
   - **Order:** BaO ¡ SrO ¡ CaO
5. **$FeCl_2$, $FeCl_3$**
   - Charges: In $FeCl_2$, Iron is Fe^{2+}. In $FeCl_3$, Iron is Fe^{3+}. The anion is Cl^- in both.
   - The magnitude of the charge on the iron cation is greater in $FeCl_3$.
   - **Order:** $FeCl_2$ ¡ $FeCl_3$

3. Predicting Ionic Formulas (Criss-Cross Method)

- To predict the formula of an ionic compound, the total positive charge from the cations must balance the total negative charge from the anions, making the compound electrically neutral.
- **Criss-Cross Method:**
  1. Write the symbols for the cation and anion, including their charges.

2. Take the numerical value of the cation's charge and use it as the subscript for the anion.
3. Take the numerical value of the anion's charge and use it as the subscript for the cation.
4. Simplify the subscripts to the lowest whole-number ratio.

Practice Problems: Write the chemical formula for the compound formed by each pair of ions.

1. **Ca^{2+} and P^{3-}**: The 2 from calcium becomes the subscript for phosphorus, and the 3 from phosphorus becomes the subscript for calcium. **Formula: Ca_3P_2**

2. **Al^{3+} and $SO_4{}^{2-}$**: The 3 from aluminum becomes the subscript for the entire sulfate ion (use parentheses), and the 2 from sulfate becomes the subscript for aluminum. **Formula: $Al_2(SO_4)_3$**

3. **Mg^{2+} and O^{2-}**: Criss-crossing gives Mg_2O_2. Ionic formulas must be simplified to the lowest whole-number ratio, which is 1:1. **Formula: MgO**

4. **$NH_4{}^{+}$ and $CO_3{}^{2-}$**: The 1 from ammonium becomes the subscript for carbonate, and the 2 from carbonate becomes the subscript for the entire ammonium ion (use parentheses). **Formula: $(NH_4)_2CO_3$**

5. **Lead(IV) and Oxide (Pb^{4+} and O^{2-})**: Criss-crossing gives Pb_2O_4. This simplifies to the lowest ratio. **Formula: PbO_2**

4. Memorization: Expected Charges of Main-Group Elements

You must memorize the typical charges for elements in the following groups:

- **Group 1 (Alkali Metals):** +1 (e.g., Li^+, Na^+, K^+)
- **Group 2 (Alkaline Earth Metals):** +2 (e.g., Mg^{2+}, Ca^{2+}, Sr^{2+})
- **Group 13:** +3 (e.g., Al^{3+})
- **Group 15:** -3 (e.g., N^{3-}, P^{3-})
- **Group 16:** -2 (e.g., O^{2-}, S^{2-})
- **Group 17 (Halogens):** -1 (e.g., F^-, Cl^-, Br^-)
- **Transition Metals:** Have variable charges and require Roman numerals in their names (e.g., Iron can be Fe^{2+} or Fe^{3+}). Zinc (Zn^{2+}), Silver (Ag^+), and Cadmium (Cd^{2+}) are exceptions that typically only have one charge.

5. Memorization: Polyatomic Ions

You must memorize the names, formulas, and charges of the following common polyatomic ions. Flashcards are highly recommended.

- **Acetate:** $C_2H_3O_2{}^-$
- **Carbonate:** $CO_3{}^{2-}$
- **Hydrogen Carbonate (Bicarbonate):** $HCO_3{}^-$
- **Hydroxide:** OH^-
- **Nitrite:** $NO_2{}^-$

- **Nitrate:** $NO_3{}^-$
- **Chromate:** $CrO_4{}^{2-}$
- **Dichromate:** $Cr_2O_7{}^{2-}$
- **Phosphate:** $PO_4{}^{3-}$
- **Hydrogen Phosphate:** $HPO_4{}^{2-}$
- **Dihydrogen Phosphate:** $H_2PO_4{}^-$
- **Ammonium:** $NH_4{}^+$
- **Hypochlorite:** ClO^-
- **Chlorite:** $ClO_2{}^-$
- **Chlorate:** $ClO_3{}^-$
- **Perchlorate:** $ClO_4{}^-$
- **Permanganate:** $MnO_4{}^-$
- **Sulfite:** $SO_3{}^{2-}$
- **Sulfate:** $SO_4{}^{2-}$
- **Hydrogen Sulfite (Bisulfite):** $HSO_3{}^-$
- **Hydrogen Sulfate (Bisulfate):** $HSO_4{}^-$
- **Cyanide:** CN^-
- **Peroxide:** $O_2{}^{2-}$

6. Chemical Nomenclature

Nomenclature is a critical skill. It will be a significant portion of the exam.

Ionic Nomenclature

- **Rules:** Name the cation (metal) first. If it's a main-group metal with a fixed charge, just use the element name. If it's a transition metal (or post-transition) with variable charges, use a Roman numeral in parentheses to indicate the charge. Then, name the anion. For a monatomic anion, change the element's ending to "-ide". For a polyatomic anion, use its memorized name.

Practice Problems (Formula to Name):

1. $Ca(NO_3)_2$: Calcium is a group 2 metal (fixed charge +2). $NO_3{}^-$ is nitrate. **Name: Calcium nitrate**
2. Fe_2O_3: Oxygen (oxide) has a -2 charge. Three oxides give a total charge of -6. To balance this, two iron atoms must have a total charge of +6, meaning each is +3. Iron is a transition metal, so use a Roman numeral. **Name: Iron(III) oxide**
3. NH_4Cl: $NH_4{}^+$ is the ammonium ion. Cl is chlorine, which becomes chloride as an anion. **Name: Ammonium chloride**
4. SnS_2: Sulfur (sulfide) has a -2 charge. Two sulfides give a total of -4. The single tin (Sn) atom must have a +4 charge. Tin is a post-transition metal with variable charge. **Name: Tin(IV) sulfide**

5. K_2SO_3: Potassium is a group 1 metal (fixed charge +1). ${SO_3}^{2-}$ is the sulfite ion. **Name: Potassium sulfite**

Practice Problems (Name to Formula):

1. **Magnesium hydroxide**: Magnesium is Mg^{2+}. Hydroxide is OH^-. Criss-cross gives $Mg_1(OH)_2$. **Formula: $Mg(OH)_2$**

2. **Copper(II) phosphate**: Copper(II) is Cu^{2+}. Phosphate is ${PO_4}^{3-}$. Criss-cross gives $Cu_3(PO_4)_2$. **Formula: $Cu_3(PO_4)_2$**

3. **Aluminum sulfide**: Aluminum is Al^{3+}. Sulfide is S^{2-}. Criss-cross gives Al_2S_3. **Formula: Al_2S_3**

4. **Lead(IV) sulfate**: Lead(IV) is Pb^{4+}. Sulfate is ${SO_4}^{2-}$. Criss-cross gives $Pb_2(SO_4)_4$, which simplifies to $Pb(SO_4)_2$. **Formula: $Pb(SO_4)_2$**

5. **Ammonium dichromate**: Ammonium is ${NH_4}^+$. Dichromate is ${Cr_2O_7}^{2-}$. Criss-cross gives $(NH_4)_2Cr_2O_7$. **Formula: $(NH_4)_2Cr_2O_7$**

Molecular Nomenclature

- **Rules:** Name the first nonmetal using its element name. Name the second nonmetal by changing its ending to "-ide". Use Greek prefixes (mono-, di-, tri-, tetra-, penta-, etc.) to indicate the number of atoms of each element. The prefix "mono-" is usually omitted for the first element.

Practice Problems (Formula to Name):

1. P_4O_{10}: Four phosphorus, ten oxygen. **Name: Tetraphosphorus decoxide**

2. SF_6: One sulfur, six fluorine. **Name: Sulfur hexafluoride**

3. N_2O_3: Two nitrogen, three oxygen. **Name: Dinitrogen trioxide**

4. IF_5: One iodine, five fluorine. **Name: Iodine pentafluoride**

5. CO: One carbon, one oxygen. **Name: Carbon monoxide**

Practice Problems (Name to Formula):

1. **Dichlorine heptoxide**: Two chlorine (Cl_2), seven oxygen (O_7). **Formula: Cl_2O_7**

2. **Carbon tetrachloride**: One carbon (C), four chlorine (Cl_4). **Formula: CCl_4**

3. **Disulfur dibromide**: Two sulfur (S_2), two bromine (Br_2). **Formula: S_2Br_2**

4. **Xenon tetrafluoride**: One xenon (Xe), four fluorine (F_4). **Formula: XeF_4**

5. **Phosphorus trichloride**: One phosphorus (P), three chlorine (Cl_3). **Formula: PCl_3**

Acid Nomenclature

- **Binary Acids:** Compounds of H with a nonmetal (usually a halogen) in aqueous solution. **Rule:** ‘hydro-‘ + base name of nonmetal + ‘-ic acid‘.

- **Oxyacids:** Compounds of H, O, and another nonmetal. **Rule:** Identify the polyatomic anion. If the anion name ends in ‘-ate‘, change it to ‘-ic acid‘. If the anion name ends in ‘-ite‘, change it to ‘-ous acid‘. Prefixes like ‘per-‘ and ‘hypo-‘ are retained.

Practice Problems (Mixed):

1. Name $H_2S(aq)$: It is H and a nonmetal. **Name: Hydrosulfuric acid**

2. Give the formula for Phosphoric acid: The '-ic' ending means it came from the 'phosphate' ion ($PO_4{}^{3-}$). Add H+ ions to balance the charge. **Formula: $\mathbf{H_3PO_4}$**

3. Name HNO_2(aq): The anion is $NO_2{}^{-}$, which is 'nitrite'. The '-ite' ending changes to '-ous acid'. **Name: Nitrous acid**

4. Give the formula for Perchloric acid: The 'per-' and '-ic' ending means it came from the 'perchlorate' ion ($ClO_4{}^{-}$). Add one H+ to balance the charge. **Formula: $\mathbf{HClO_4}$**

5. Name HCN(aq): While not strictly binary, it is named like one. The anion is cyanide. **Name: Hydrocyanic acid**

Hydrate Nomenclature

- **Rule:** Name the ionic compound normally, then add a Greek prefix to indicate the number of water molecules, followed by the word "hydrate".

Practice Problems (Mixed):

1. Name $CuSO_4 \cdot 5\,H_2O$: Copper can have multiple charges. Sulfate ($SO_4{}^{2-}$) has a -2 charge, so the copper must be +2. There are five water molecules. **Name: Copper(II) sulfate pentahydrate**

2. Give the formula for Barium chloride dihydrate: Barium is Ba^{2+}. Chloride is Cl^{-}. The ionic compound is $BaCl_2$. "Dihydrate" means two water molecules. **Formula: $\mathbf{BaCl_2 \cdot 2\,H_2O}$**

3. Name $FePO_4 \cdot 4\,H_2O$: Phosphate ($PO_4{}^{3-}$) has a -3 charge, so the iron must be +3. There are four water molecules. **Name: Iron(III) phosphate tetrahydrate**

4. Give the formula for Sodium carbonate decahydrate: Sodium is Na^{+}. Carbonate is $CO_3{}^{2-}$. The compound is Na_2CO_3. "Decahydrate" means ten waters. **Formula: $\mathbf{Na_2CO_3 \cdot 10\,H_2O}$**

5. Name $MgSO_4 \cdot 7\,H_2O$: Magnesium is Mg^{2+}. Sulfate is $SO_4{}^{2-}$. The compound is MgSO4. "Heptahydrate" for seven waters. **Name: Magnesium sulfate heptahydrate**

7. Empirical and Molecular Formulas

- **Empirical Formula:** The simplest whole-number ratio of atoms in a compound.
- **Molecular Formula:** The actual number of atoms of each element in one molecule of the compound. It is a whole-number multiple of the empirical formula.

Practice Problems: Find the empirical formula for each molecular formula.

1. **$\mathbf{C_6H_{12}O_6}$ (Glucose):** All subscripts (6, 12, 6) are divisible by 6. Dividing gives a 1:2:1 ratio. **Empirical Formula: $\mathbf{CH_2O}$**

2. **$\mathbf{N_2O_4}$ (Dinitrogen tetroxide):** Both subscripts (2, 4) are divisible by 2. Dividing gives a 1:2 ratio. **Empirical Formula: $\mathbf{NO_2}$**

3. **$\mathbf{C_5H_{12}}$ (Pentane):** The subscripts (5, 12) have no common divisor other than 1. **Empirical Formula: $\mathbf{C_5H_{12}}$**

4. **$\mathbf{H_2O_2}$ (Hydrogen peroxide):** Both subscripts (2, 2) are divisible by 2. Dividing gives a 1:1 ratio. **Empirical Formula: HO**

5. **$\mathbf{C_8H_{10}N_4O_2}$ (Caffeine):** All subscripts (8, 10, 4, 2) are divisible by 2. Dividing gives a 4:5:2:1 ratio. **Empirical Formula: $\mathbf{C_4H_5N_2O}$**

8. Molar Mass and Mass Percent

- **Molar Mass:** The mass in grams of one mole of a substance. It is calculated by summing the atomic masses of all atoms in the chemical formula.
- **Mass Percent Composition:** The percentage by mass of each element in a compound.

$$\text{Mass \% of element} = \frac{(\text{number of atoms of element}) \times (\text{atomic mass of element})}{\text{molar mass of compound}} \times 100\%$$

Practice Problems (Molar Mass):

1. $\mathbf{H_2SO_4}$: $2(1.01) + 32.07 + 4(16.00) = \mathbf{98.09\ g/mol}$
2. $\mathbf{Ca(NO_3)_2}$: $40.08 + 2(14.01) + 6(16.00) = \mathbf{164.10\ g/mol}$
3. $\mathbf{C_{12}H_{22}O_{11}}$: $12(12.01) + 22(1.01) + 11(16.00) = \mathbf{342.34\ g/mol}$
4. $\mathbf{Fe_2(SO_4)_3}$: $2(55.85) + 3(32.07) + 12(16.00) = \mathbf{399.91\ g/mol}$
5. $\mathbf{Mg(OH)_2}$: $24.31 + 2(16.00) + 2(1.01) = \mathbf{58.33\ g/mol}$

Practice Problems (Mass Percent):

1. **Find the mass % of C in $\mathbf{C_3H_8}$ (Propane). Molar mass = 44.11 g/mol**:

$$\frac{3 \times 12.01\text{ g/mol}}{44.11\text{ g/mol}} \times 100\% = \mathbf{81.68\%\ C}$$

2. **Find the mass % of O in $\mathbf{H_2SO_4}$. Molar mass = 98.09 g/mol**:

$$\frac{4 \times 16.00\text{ g/mol}}{98.09\text{ g/mol}} \times 100\% = \mathbf{65.25\%\ O}$$

3. **Find the mass % of N in Ammonium Nitrate ($\mathbf{NH_4NO_3}$). Molar mass = 80.06 g/mol**:

$$\frac{2 \times 14.01\text{ g/mol}}{80.06\text{ g/mol}} \times 100\% = \mathbf{35.00\%\ N}$$

4. **Find the mass % of $\mathbf{H_2O}$ in Copper(II) Sulfate Pentahydrate ($\mathbf{CuSO_4 \cdot 5\,H_2O}$). Molar mass = 249.72 g/mol**:

$$\frac{5 \times 18.02\text{ g/mol}}{249.72\text{ g/mol}} \times 100\% = \mathbf{36.08\%\ H_2O}$$

5. **Find the mass % of each element in $\mathbf{CaCO_3}$. Molar mass = 100.09 g/mol**:
   - % Ca: $\frac{40.08}{100.09} \times 100\% = \mathbf{40.04\%}$
   - % C: $\frac{12.01}{100.09} \times 100\% = \mathbf{12.00\%}$
   - % O: $\frac{3\times16.00}{100.09} \times 100\% = \mathbf{47.96\%}$

9. Conversions: Grams, Moles, and Atoms/Molecules

- Use **Molar Mass** to convert between grams and moles.
- Use **Avogadro's Number** (6.022×10^{23}) to convert between moles and atoms/molecules.

Practice Problems:

1. **How many moles are in 50.0 g of $CaCO_3$ (molar mass = 100.09 g/mol)?**

$$50.0 \text{ g CaCO}_3 \times \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} = \mathbf{0.500 \text{ mol CaCO}_3}$$

2. **What is the mass in grams of 2.50 moles of H_2O (molar mass = 18.02 g/mol)?**

$$2.50 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \mathbf{45.1 \text{ g H}_2\text{O}}$$

3. **How many molecules are in 3.00 moles of CO_2?**

$$3.00 \text{ mol CO}_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \mathbf{1.81 \times 10^{24} \text{ molecules CO}_2}$$

4. **How many atoms of oxygen are in 22.0 g of CO_2 (molar mass = 44.01 g/mol)?**

$$22.0 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules CO}_2}{1 \text{ mol CO}_2} \times \frac{2 \text{ atoms O}}{1 \text{ molecule CO}_2} = \mathbf{6.02 \times 10^{23} \text{ atoms O}}$$

5. **What is the mass in grams of a single molecule of H_2SO_4 (molar mass = 98.09 g/mol)?**

$$1 \text{ molecule} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{98.09 \text{ g}}{1 \text{ mol}} = \mathbf{1.63 \times 10^{-22} \text{ g}}$$

Chapter 8: Chemical Reactions

1. Balancing Chemical Equations

- The Law of Conservation of Mass dictates that atoms are neither created nor destroyed in a chemical reaction. Therefore, the number of atoms of each element must be the same on both the reactant and product sides of an equation.
- Balancing is achieved by placing **stoichiometric coefficients** in front of the chemical formulas. **Never change the subscripts!**

Practice Problems: Balance the following equations.

1. **_ C_3H_8 + _ O_2 → _ CO_2 + _ H_2O**
   - Balance C: $C_3H_8 + O_2 \longrightarrow 3\,CO_2 + H_2O$
   - Balance H: $C_3H_8 + O_2 \longrightarrow 3\,CO_2 + 4\,H_2O$
   - Balance O: There are $3 \times 2 + 4 \times 1 = 10$ oxygens on the right. Need 5 O_2 on the left.
   - **Answer:** $C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O$
2. **_ Fe + _ H_2O → _ Fe_3O_4 + _ H_2**
   - Balance Fe: $3\,Fe + H_2O \longrightarrow Fe_3O_4 + H_2$

- Balance O: $3\,Fe + 4\,H_2O \longrightarrow Fe_3O_4 + H_2$
- Balance H: $3\,Fe + 4\,H_2O \longrightarrow Fe_3O_4 + 4\,H_2$
- **Answer:** $3\,Fe + 4\,H_2O \longrightarrow Fe_3O_4 + 4\,H_2$

3. **_ Al + _ H_2SO_4 → _ $Al_2(SO_4)_3$ + _ H_2**
   - Balance Al: $2\,Al + H_2SO_4 \longrightarrow Al_2(SO_4)_3 + H_2$
   - Balance the sulfate group (SO_4) as a whole: $2\,Al + 3\,H_2SO_4 \longrightarrow Al_2(SO_4)_3 + H_2$
   - Balance H: $2\,Al + 3\,H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3\,H_2$
   - **Answer:** $2\,Al + 3\,H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3\,H_2$

4. **_ C_2H_6 + _ O_2 → _ CO_2 + _ H_2O**
   - Balance C: $C_2H_6 + O_2 \longrightarrow 2\,CO_2 + H_2O$
   - Balance H: $C_2H_6 + O_2 \longrightarrow 2\,CO_2 + 3\,H_2O$
   - Balance O: Right side has $2 \times 2 + 3 \times 1 = 7$ oxygens. To get 7 on the left, we need 7/2 O_2.
   - $C_2H_6 + \frac{7}{2}\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O$. Multiply everything by 2 to clear the fraction.
   - **Answer:** $2\,C_2H_6 + 7\,O_2 \longrightarrow 4\,CO_2 + 6\,H_2O$

5. **_ NH_3 + _ CuO → _ Cu + _ N_2 + _ H_2O**
   - Balance N: $2\,NH_3 + CuO \longrightarrow Cu + N_2 + H_2O$
   - Balance H: $2\,NH_3 + CuO \longrightarrow Cu + N_2 + 3\,H_2O$
   - Balance O: $2\,NH_3 + 3\,CuO \longrightarrow Cu + N_2 + 3\,H_2O$
   - Balance Cu: $2\,NH_3 + 3\,CuO \longrightarrow 3\,Cu + N_2 + 3\,H_2O$
   - **Answer:** $2\,NH_3 + 3\,CuO \longrightarrow 3\,Cu + N_2 + 3\,H_2O$

2. Classifying Reaction Types

- **Synthesis (Combination):** Two or more reactants combine to form a single product. (A + B → AB)
- **Decomposition:** A single compound breaks down into two or more simpler substances. (AB → A + B)
- **Single Replacement:** An element reacts with a compound, displacing another element from it. (A + BC → AC + B)
- **Double Replacement:** The cations of two ionic compounds exchange places. (AB + CD → AD + CB)
- **Hydrocarbon Combustion:** A hydrocarbon (C_xH_y) or ($C_xH_yO_z$) reacts with oxygen (O_2) to produce carbon dioxide (CO_2) and water (H_2O).

Practice Problems: Classify each reaction.

1. $2\,H_2 + O_2 \longrightarrow 2\,H_2O$: Two elements form one compound. **Synthesis**
2. $Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$: Mg displaces H. **Single Replacement**
3. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$: One compound breaks down. **Decomposition**
4. $C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g)$: Hydrocarbon + O2 yields CO2 + H2O. **Hydrocarbon Combustion**
5. $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$: Ag and Na swap places. **Double Replacement**

3. **Combustion Analysis, Empirical & Molecular Formulas**

- **Combustion Analysis** is a technique used to determine the empirical formula of a compound. The compound is burned in excess oxygen, and the masses of CO_2 and H_2O produced are measured.

- **Calculation Steps:**

  1. Convert mass of CO_2 to moles of CO_2, then to moles of C.
  2. Convert mass of H_2O to moles of H_2O, then to moles of H.
  3. If the compound contains oxygen, find the mass of C and H, subtract from the total sample mass to find the mass of O, then convert mass of O to moles of O.
  4. Divide all mole values by the smallest mole value to get the empirical formula ratio. If necessary, multiply by a small integer to get whole numbers (e.g., if you get X.5, multiply by 2; if you get X.33 or X.67, multiply by 3).
  5. To find the molecular formula, calculate the molar mass of the empirical formula. Divide the given molecular mass by the empirical formula mass to get a whole number, n. Multiply the subscripts in the empirical formula by n.

Practice Problems:

1. **A 1.50 g sample of a hydrocarbon undergoes combustion to produce 4.40 g of $\mathbf{CO_2}$ and 2.70 g of $\mathbf{H_2O}$. Determine its empirical formula.**

$$\text{mol C} = 4.40 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g}} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.100 \text{ mol C}$$
$$\text{mol H} = 2.70 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g}} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.300 \text{ mol H}$$

Ratio C:H is 0.100:0.300. Divide by smallest (0.100) → 1:3. **Empirical Formula: $\mathbf{CH_3}$**

2. **If the molar mass of the compound in the problem above is 30.07 g/mol, what is its molecular formula?**

   - Empirical formula mass of $CH_3 = 12.01 + 3(1.01) = 15.04$ g/mol.
   - $n = \frac{\text{Molecular Mass}}{\text{Empirical Mass}} = \frac{30.07}{15.04} \approx 2$.
   - Molecular formula $= (CH_3)_2$. **Molecular Formula: $\mathbf{C_2H_6}$**

3. **Combustion of 0.8233 g of a compound containing C, H, and O produces 2.445 g of $\mathbf{CO_2}$ and 0.6003 g of $\mathbf{H_2O}$. Determine the empirical formula.**

$$\text{g C} = 2.445 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 0.6672 \text{ g C}$$
$$\text{g H} = 0.6003 \text{ g } H_2O \times \frac{2.02 \text{ g H}}{18.02 \text{ g } H_2O} = 0.0671 \text{ g H}$$
$$\text{g O} = 0.8233 \text{ g sample} - (0.6672 \text{ g C} + 0.0671 \text{ g H}) = 0.0890 \text{ g O}$$

Now, convert grams to moles:

$$\text{mol C} = 0.6672 \text{ g C}/12.01 \text{ g/mol} = 0.0556 \text{ mol C}$$
$$\text{mol H} = 0.0671 \text{ g H}/1.01 \text{ g/mol} = 0.0664 \text{ mol H}$$
$$\text{mol O} = 0.0890 \text{ g O}/16.00 \text{ g/mol} = 0.00556 \text{ mol O}$$

Divide by smallest (0.00556): C: 10, H: 11.9 ≈ 12, O: 1. **Empirical Formula: $\mathbf{C_{10}H_{12}O}$**

4. **A 0.250 g sample of a compound containing C, H, and O is burned. The reaction produces 0.366 g of $\mathbf{CO_2}$ and 0.150 g of $\mathbf{H_2O}$. Its molar mass is 60.1 g/mol. Find its molecular formula.**

$$\text{g C} = 0.366 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 0.100 \text{ g C} \rightarrow 0.00833 \text{ mol C}$$
$$\text{g H} = 0.150 \text{ g } H_2O \times \frac{2.02 \text{ g H}}{18.02 \text{ g } H_2O} = 0.0168 \text{ g H} \rightarrow 0.0166 \text{ mol H}$$
$$\text{g O} = 0.250 - (0.100 + 0.0168) = 0.1332 \text{ g O} \rightarrow 0.00833 \text{ mol O}$$

Divide by smallest (0.00833): C: 1, H: 2, O: 1. Empirical formula is CH_2O. Empirical mass = 30.03 g/mol. $n = 60.1/30.03 \approx 2$. **Molecular Formula: $\mathbf{C_2H_4O_2}$**

5. **Ascorbic acid (Vitamin C) contains C, H, and O. Combustion of a 1.000 g sample produces 1.50 g of $\mathbf{CO_2}$ and 0.408 g of $\mathbf{H_2O}$. Determine the empirical and molecular formula, given the molar mass is 176 g/mol.**

$$\text{g C} = 1.50 \text{ g } CO_2 \times \frac{12.01}{44.01} = 0.409 \text{ g C} \rightarrow 0.0341 \text{ mol C}$$
$$\text{g H} = 0.408 \text{ g } H_2O \times \frac{2.02}{18.02} = 0.0456 \text{ g H} \rightarrow 0.0452 \text{ mol H}$$
$$\text{g O} = 1.000 - (0.409 + 0.0456) = 0.545 \text{ g O} \rightarrow 0.0341 \text{ mol O}$$

Divide by smallest (0.0341): C: 1, H: 1.33, O: 1. To get whole numbers, multiply by 3. Ratio is 3:4:3. Empirical formula is $C_3H_4O_3$. Empirical mass = 88.06 g/mol. $n = 176/88.06 \approx 2$. **Molecular Formula: $\mathbf{C_6H_8O_6}$**

4. Stoichiometry, Limiting Reactants, and Yield

- **Stoichiometry** uses mole ratios from a balanced chemical equation to relate the amounts of reactants and products.
- **Limiting Reactant:** The reactant that is completely consumed in a reaction and determines the maximum amount of product that can be formed.
- **Excess Reactant:** The reactant that is left over after the reaction is complete.
- **Theoretical Yield:** The maximum amount of product that can be formed from the given amounts of reactants, calculated based on the limiting reactant.
- **Actual Yield:** The amount of product actually obtained from a reaction in the lab.
- **Percent Yield:** $\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$

Practice Problems:

1. **For the reaction $\mathbf{N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)}$, if you have 2.0 mol of $\mathbf{N_2}$ and 3.0 mol of $\mathbf{H_2}$, which is the limiting reactant and what is the theoretical yield of $\mathbf{NH_3}$ in moles?**
   - From N_2: 2.0 mol $N_2 \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 4.0$ mol NH_3
   - From H_2: 3.0 mol $H_2 \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} = 2.0$ mol NH_3
   - H_2 produces less product, so it is the **limiting reactant**.
   - The theoretical yield is **2.0 mol $\mathbf{NH_3}$**.

2. **Using the same reaction, what mass of NH_3 can be formed from 28.0 g of N_2 and 9.0 g of H_2?**

$$\text{g NH}_3 \text{ from N}_2 = 28.0 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g}} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} \times \frac{17.04 \text{ g NH}_3}{1 \text{ mol}} = 34.0 \text{ g NH}_3$$
$$\text{g NH}_3 \text{ from H}_2 = 9.0 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.02 \text{ g}} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \times \frac{17.04 \text{ g NH}_3}{1 \text{ mol}} = 50.6 \text{ g NH}_3$$

N_2 is the limiting reactant. The theoretical yield is **34.0 g NH_3**.

3. **If 15.0 g of copper(II) chloride ($CuCl_2$) reacts with 20.0 g of sodium nitrate ($NaNO_3$), how many grams of sodium chloride (NaCl) can be formed? $CuCl_2 + 2\,NaNO_3 \longrightarrow Cu(NO_3)_2 + 2\,NaCl$**

$$\text{g NaCl from CuCl}_2 = 15.0 \text{ g CuCl}_2 \times \frac{1 \text{ mol}}{134.45 \text{ g}} \times \frac{2 \text{ mol NaCl}}{1 \text{ mol CuCl}_2} \times \frac{58.44 \text{ g}}{1 \text{ mol}} = 13.0 \text{ g NaCl}$$
$$\text{g NaCl from NaNO}_3 = 20.0 \text{ g NaNO}_3 \times \frac{1 \text{ mol}}{85.00 \text{ g}} \times \frac{2 \text{ mol NaCl}}{2 \text{ mol NaNO}_3} \times \frac{58.44 \text{ g}}{1 \text{ mol}} = 13.8 \text{ g NaCl}$$

$CuCl_2$ is the limiting reactant. The theoretical yield is **13.0 g NaCl**.

4. **For the problem above, how much of the excess reagent ($NaNO_3$) is left over?**

   - First, find how much $NaNO_3$ was used:

$$15.0 \text{ g CuCl}_2 \times \frac{1 \text{ mol}}{134.45 \text{ g}} \times \frac{2 \text{ mol NaNO}_3}{1 \text{ mol CuCl}_2} \times \frac{85.00 \text{ g}}{1 \text{ mol}} = 18.9 \text{ g NaNO}_3 \text{ used}$$

   - Subtract used amount from the starting amount: 20.0 g − 18.9 g = **1.1 g left over**.

5. **In a reaction, the theoretical yield of a product is 8.50 g. The actual yield obtained in the lab was 7.82 g. What is the percent yield?**

$$\% \text{ Yield} = \frac{7.82 \text{ g}}{8.50 \text{ g}} \times 100\% = \mathbf{92.0\%}$$

Chapter 9: Reactions in Aqueous Solutions

1. Molarity Calculations

- **Molarity (M)** is a unit of concentration defined as moles of solute per liter of solution.

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{Liters of solution}}$$

Practice Problems:

1. **Calculate the molarity of a solution made by dissolving 23.4 g of NaCl in enough water to make 500. mL of solution.**

$$\text{moles NaCl} = 23.4 \text{ g} \times \frac{1 \text{ mol}}{58.44 \text{ g}} = 0.400 \text{ mol}$$
$$\text{Molarity} = \frac{0.400 \text{ mol}}{0.500 \text{ L}} = \mathbf{0.800 \text{ M}}$$

2. **How many moles of HCl are present in 25.0 mL of a 0.550 M HCl solution?**

$$\text{moles HCl} = \text{M} \times \text{L} = (0.550 \text{ mol/L}) \times (0.0250 \text{ L}) = \mathbf{0.0138 \text{ mol}}$$

3. **What volume (in mL) of a 1.50 M KNO_3 solution contains 5.00 g of KNO_3?**

$$\text{moles KNO}_3 = 5.00\text{ g} \times \frac{1\text{ mol}}{101.11\text{ g}} = 0.04945\text{ mol}$$
$$\text{Liters} = \frac{\text{moles}}{\text{M}} = \frac{0.04945\text{ mol}}{1.50\text{ mol/L}} = 0.0330\text{ L} = \mathbf{33.0\ mL}$$

4. **How many grams of NaOH are needed to prepare 250. mL of a 0.200 M solution?**

$$\text{moles NaOH} = 0.200\text{ mol/L} \times 0.250\text{ L} = 0.0500\text{ mol}$$
$$\text{grams NaOH} = 0.0500\text{ mol} \times 40.00\text{ g/mol} = \mathbf{2.00\ g}$$

5. **What is the molar concentration of chloride ions in a solution prepared by dissolving 12.5 g of $AlCl_3$ in 250 mL of water?**

$$\text{Molarity of AlCl}_3 = \frac{12.5\text{ g}/133.33\text{ g/mol}}{0.250\text{ L}} = \frac{0.09375\text{ mol}}{0.250\text{ L}} = 0.375\text{ M AlCl}_3$$
$$\text{Since AlCl}_3 \rightarrow \text{Al}^{3+} + 3\,\text{Cl}^-,\ [\text{Cl}^-] = 3 \times 0.375\text{ M} = \mathbf{1.13\ M}$$

2. Dilutions

- **Dilution** is the process of decreasing the concentration of a stock solution by adding more solvent. The amount of solute remains constant.
- **Dilution Formula:** $M_1V_1 = M_2V_2$, where M_1 and V_1 are the molarity and volume of the concentrated solution, and M_2 and V_2 are the molarity and volume of the diluted solution.

Practice Problems:

1. **What volume of 12.0 M HCl is needed to prepare 250. mL of 1.50 M HCl?**

$$M_1V_1 = M_2V_2 \rightarrow (12.0\text{ M})V_1 = (1.50\text{ M})(250.\text{ mL})$$
$$V_1 = \frac{(1.50)(250.)}{12.0} = \mathbf{31.3\ mL}$$

2. **If 50.0 mL of a 2.50 M H_2SO_4 solution is diluted to a final volume of 500. mL, what is the new concentration?**

$$(2.50\text{ M})(50.0\text{ mL}) = M_2(500.\text{ mL})$$
$$M_2 = \frac{(2.50)(50.0)}{500.} = \mathbf{0.250\ M}$$

3. **How would you prepare 100. mL of 0.400 M $MgSO_4$ from a stock solution of 2.00 M $MgSO_4$?**

$$(2.00\text{ M})V_1 = (0.400\text{ M})(100.\text{ mL}) \rightarrow V_1 = 20.0\text{ mL}$$

Answer: You would take 20.0 mL of the 2.00 M stock solution and add enough water to make a total volume of 100. mL.

4. **To what volume must you dilute 75.0 mL of a 10.0 M NaOH solution to obtain a 1.25 M solution?**

$$(10.0\text{ M})(75.0\text{ mL}) = (1.25\text{ M})V_2$$
$$V_2 = \frac{(10.0)(75.0)}{1.25} = \mathbf{600.\ mL}$$

5. **If 200. mL of water is added to 50.0 mL of 4.00 M KCl, what is the final molarity? (Assume volumes are additive).**
   - $V_2 = V_1 + \text{added water} = 50.0\text{ mL} + 200.\text{ mL} = 250.\text{ mL}$
   - $(4.00\text{ M})(50.0\text{ mL}) = M_2(250.\text{ mL}) \rightarrow M_2 = \mathbf{0.800\ M}$

3. Solution Stoichiometry (Titrations)

- **Titration** is a lab technique used to determine the concentration of an unknown solution (analyte) by reacting it with a solution of known concentration (titrant).
- The key is to use the mole ratio from the balanced chemical equation to relate the moles of titrant to the moles of analyte.
- **Path:** Volume of Titrant $\xrightarrow{\text{Molarity of Titrant}}$ Moles of Titrant $\xrightarrow{\text{Mole Ratio}}$ Moles of Analyte $\xrightarrow{\text{Volume of Analyte}}$ Molarity of Analyte.

Practice Problems:

1. **What volume of 0.125 M NaOH is required to neutralize 25.0 mL of 0.100 M HCl? $\mathbf{HCl + NaOH \longrightarrow NaCl + H_2O}$**

$$\text{mol HCl} = 0.100\text{ M} \times 0.0250\text{ L} = 0.00250\text{ mol HCl}$$
$$\text{mol NaOH} = 0.00250\text{ mol HCl} \times \frac{1\text{ mol NaOH}}{1\text{ mol HCl}} = 0.00250\text{ mol NaOH}$$
$$\text{L NaOH} = \frac{0.00250\text{ mol}}{0.125\text{ M}} = 0.0200\text{ L} = \mathbf{20.0\ mL}$$

2. **If 35.45 mL of 0.175 M $\mathbf{H_2SO_4}$ is required to neutralize 25.00 mL of a KOH solution, what is the molarity of the KOH solution? $\mathbf{H_2SO_4 + 2\,KOH \longrightarrow K_2SO_4 + 2\,H_2O}$**

$$\text{mol } H_2SO_4 = 0.175\text{ M} \times 0.03545\text{ L} = 0.006204\text{ mol } H_2SO_4$$
$$\text{mol KOH} = 0.006204\text{ mol } H_2SO_4 \times \frac{2\text{ mol KOH}}{1\text{ mol } H_2SO_4} = 0.01241\text{ mol KOH}$$
$$\text{Molarity KOH} = \frac{0.01241\text{ mol}}{0.02500\text{ L}} = \mathbf{0.496\ M}$$

3. **What is the molarity of a $\mathbf{Ba(OH)_2}$ solution if 44.18 mL is needed to neutralize 20.00 mL of 0.1016 M HCl?**
   - Balanced Equation: $2\,HCl + Ba(OH)_2 \longrightarrow BaCl_2 + 2\,H_2O$
   - Moles HCl: $0.1016\text{ M} \times 0.02000\text{ L} = 0.002032\text{ mol HCl}$
   - Moles $Ba(OH)_2$: $0.002032\text{ mol HCl} \times \frac{1\text{ mol } Ba(OH)_2}{2\text{ mol HCl}} = 0.001016\text{ mol } Ba(OH)_2$
   - Molarity $Ba(OH)_2$: $\frac{0.001016\text{ mol}}{0.04418\text{ L}} = \mathbf{0.0230\ M}$

4. **A 0.552 g sample of KHP ($\mathbf{KHC_8H_4O_4}$, molar mass = 204.22 g/mol) is titrated with a NaOH solution. If 23.45 mL of the base is required, what is the base's molarity? (KHP is monoprotic).**
   - Balanced Equation: $KHP + NaOH \longrightarrow KNaP + H_2O$ (1:1 ratio)
   - Moles KHP: $0.552\text{ g}/204.22\text{ g/mol} = 0.002703\text{ mol KHP}$
   - Moles NaOH = Moles KHP = 0.002703 mol
   - Molarity NaOH: $\frac{0.002703\text{ mol}}{0.02345\text{ L}} = \mathbf{0.115\ M}$

5. **How many grams of $Mg(OH)_2$ would be needed to neutralize 250. mL of 0.500 M HCl?**

- Balanced Equation: $2\,HCl + Mg(OH)_2 \longrightarrow MgCl_2 + 2\,H_2O$
- Moles HCl: 0.500 M × 0.250 L = 0.125 mol HCl
- Moles $Mg(OH)_2$: 0.125 mol HCl × $\frac{1 \text{ mol } Mg(OH)_2}{2 \text{ mol HCl}}$ = 0.0625 mol $Mg(OH)_2$
- Grams $Mg(OH)_2$: 0.0625 mol × 58.33 g/mol = **3.65 g**

4. pH and $[H_3O^+]$ Calculations

- **pH** is a measure of acidity. $pH = -\log[H_3O^+]$
- **Hydronium ion concentration:** $[H_3O^+] = 10^{-pH}$
- **Scale:** Acidic: pH ¡ 7. Neutral: pH = 7. Basic: pH ¿ 7.

Practice Problems ($[H_3O^+]$ to pH):

1. $\mathbf{[H_3O^+]} = 1.0 \times 10^{-4}$ **M**: $pH = -\log(1.0 \times 10^{-4}) = 4.00$. **Acidic**
2. $\mathbf{[H_3O^+]} = 3.5 \times 10^{-9}$ **M**: $pH = -\log(3.5 \times 10^{-9}) = 8.46$. **Basic**
3. $\mathbf{[H_3O^+]} = 2.1 \times 10^{-2}$ **M**: $pH = -\log(2.1 \times 10^{-2}) = 1.68$. **Acidic**
4. $\mathbf{[H_3O^+]} = 1.0 \times 10^{-7}$ **M**: $pH = -\log(1.0 \times 10^{-7}) = 7.00$. **Neutral**
5. $\mathbf{[H_3O^+]} = 7.8 \times 10^{-12}$ **M**: $pH = -\log(7.8 \times 10^{-12}) = 11.11$. **Basic**

Practice Problems (pH to $[H_3O^+]$):

1. **pH = 3.00**: $[H_3O^+] = 10^{-3.00} =$

1. **pH = 10.50**: $[H_3O^+] = 10^{-10.50} =$

1. **pH = 6.80**: $[H_3O^+] = 10^{-6.80} =$

1. **pH = 1.25**: $[H_3O^+] = 10^{-1.25} =$

1. **pH = 8.95**: $[H_3O^+] = 10^{-8.95} =$

Topics Excluded From This Exam

Based on the professor's instructions, the following topics, while present in the course materials, are not on this specific exam and will likely be on a future exam:

- From Chapter 9, any practice documents beyond "Molarity, dilutions, and solution stoichiometry" are intended for Exam 4. This likely includes topics such as:
  - Detailed Redox Reactions (balancing in acidic/basic solution)
  - Electrolytes and Net Ionic Equations (beyond their basic application in stoichiometry)
  - Oxidation States and Redox Titrations

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