Comprehensive Study Guide for Chemistry Exam 2

Chapter 5: Ionic and Covalent Compounds

Ionic vs. Molecular Compounds

- **Ionic Compound:** A compound formed from the electrostatic attraction between a **metal cation** (positive ion) and a **nonmetal anion** (negative ion). Electrons are transferred from the metal to the nonmetal. Example: NaCl (Sodium is a metal, Chlorine is a nonmetal).
- **Molecular (Covalent) Compound:** A compound formed from the sharing of electrons between two or more **nonmetals**. Example: H_2O (Hydrogen and Oxygen are both nonmetals).

Practice: Classifying Ionic Compounds

1. **Classify $MgCl_2$:** Magnesium (Mg) is a Group 2 metal. Chlorine (Cl) is a Group 17 nonmetal. This is a metal-nonmetal combination, so it is **ionic**.
2. **Classify $FePO_4$:** Iron (Fe) is a transition metal. Phosphate (PO_4^{3-}) is a polyatomic anion. The compound consists of a metal cation and an anion, making it **ionic**.
3. **Classify K_3N:** Potassium (K) is a Group 1 metal. Nitrogen (N) is a Group 15 nonmetal. This is a metal-nonmetal combination, so it is **ionic**.
4. **Classify $CuSO_4$:** Copper (Cu) is a transition metal. Sulfate (SO_4^{2-}) is a polyatomic anion. The compound is formed between a metal cation and an anion, making it **ionic**.
5. **Classify NH_4Cl:** Ammonium (NH_4^+) is a polyatomic cation (composed of nonmetals). Chloride (Cl^-) is a nonmetal anion. Even though it's made entirely of nonmetals, the bond is between a cation and an anion, so it is classified as **ionic**.

Practice: Classifying Molecular Compounds

1. **Classify CO_2:** Carbon (C) is a nonmetal. Oxygen (O) is a nonmetal. This is a nonmetal-nonmetal combination, so it is **molecular**.
2. **Classify P_4O_{10}:** Phosphorus (P) is a nonmetal. Oxygen (O) is a nonmetal. This is a nonmetal-nonmetal combination, so it is **molecular**.
3. **Classify S_2Cl_4:** Sulfur (S) is a nonmetal. Chlorine (Cl) is a nonmetal. This is a nonmetal-nonmetal combination, so it is **molecular**.
4. **Classify NBr_3:** Nitrogen (N) is a nonmetal. Bromine (Br) is a nonmetal. This is a nonmetal-nonmetal combination, so it is **molecular**.
5. **Classify H_2S:** Hydrogen (H) is a nonmetal. Sulfur (S) is a nonmetal. This is a nonmetal-nonmetal combination, so it is **molecular**.

Lattice Energy

- **Definition:** Lattice energy is a measure of the stability of an ionic solid. It is the energy released when gaseous ions combine to form one mole of a solid ionic compound. A higher lattice energy indicates a stronger ionic bond and a more stable compound.

- **Trends:** Lattice energy is governed by Coulomb's Law: $E \propto \frac{|Q_1 \times Q_2|}{d}$
  - **Ionic Charge (Q):** Lattice energy increases dramatically as the magnitude of the ionic charges increases. This is the most important factor.
  - **Ionic Radius (distance, d):** Lattice energy decreases as the size of the ions increases (because the distance between their centers increases).

Practice Problems: Ranking by Lattice Energy

1. **Arrange in order of increasing lattice energy: KCl, SrS, RbI.**
   - **Charges:** KCl (+1,-1), SrS (+2,-2), RbI (+1,-1).
   - **Analysis:** SrS has the highest product of charges (—+2 × -2— = 4), so it will have the highest lattice energy. Both KCl and RbI have a charge product of 1. Between them, Rb^+ is larger than K^+ and I^- is larger than Cl^-. Therefore, the distance in RbI is greater, giving it a lower lattice energy than KCl.
   - **Answer: RbI ¡ KCl ¡ SrS**
2. **Arrange in order of increasing lattice energy: AlF_3, $AlCl_3$, $AlBr_3$.**
   - **Charges:** All compounds involve Al^{3+} and a halide ion (X^-). The charge factor is the same for all.
   - **Radius:** The size of the anions increases down the group: F^- ¡ Cl^- ¡ Br^-.
   - **Conclusion:** Since lattice energy is inversely proportional to ionic radius, the compound with the smallest anion (AlF_3) will have the highest lattice energy.
   - **Answer: $AlBr_3$ ¡ $AlCl_3$ ¡ AlF_3**
3. **Which has the higher lattice energy: NaF or MgO?**
   - **Charges:** NaF has (+1, -1) ions. MgO has (+2, -2) ions.
   - **Conclusion:** The product of charges for MgO (4) is much greater than for NaF (1). This charge difference is the dominant factor, making the lattice energy of **MgO** significantly higher.
4. **Arrange in order of increasing lattice energy: $MgCl_2$, $BeCl_2$, $CaCl_2$.**
   - **Charges:** All compounds contain a Group 2 cation (+2) and chloride anions (-1). The charge factor is the same.
   - **Radius:** The size of the cations increases down the group: Be^{2+} ¡ Mg^{2+} ¡ Ca^{2+}.
   - **Conclusion:** The compound with the smallest cation ($BeCl_2$) has the smallest distance between ions and thus the highest lattice energy.
   - **Answer: $CaCl_2$ ¡ $MgCl_2$ ¡ $BeCl_2$**
5. **Arrange in order of increasing lattice energy: NaCl, MgS, AlP.**
   - **Charges:** NaCl (+1, -1), MgS (+2, -2), AlP (+3, -3).
   - **Conclusion:** The product of the ionic charges increases significantly (1, 4, 9). This is the dominant factor.
   - **Answer: NaCl ¡ MgS ¡ AlP**

Predicting Ionic Formulas (Criss-Cross Method)

To Do: Predict the chemical formula of an ionic compound by balancing the charges. The total positive charge from the cations must equal the total negative charge from the anions. The "criss-cross" method is a shortcut: the magnitude of the charge on one ion becomes the subscript for the other ion. Always simplify the subscripts to the smallest whole-number ratio.

Practice Problems: Predicting Formulas

1. **Formula for Aluminum Oxide:** Aluminum (Group 13) is Al^{3+}. Oxide (Group 16) is O^{2-}. Criss-crossing the charges gives $\mathbf{Al_2O_3}$.

2. **Formula for Magnesium Nitride:** Magnesium (Group 2) is Mg^{2+}. Nitride (Group 15) is N^{3-}. Criss-crossing the charges gives $\mathbf{Mg_3N_2}$.

3. **Formula for Scandium (III) Hydroxide:** Scandium (III) is Sc^{3+}. Hydroxide is the polyatomic ion OH^-. Criss-crossing gives $Sc(OH)_3$. Parentheses are required around the polyatomic ion. The formula is $\mathbf{Sc(OH)_3}$.

4. **Formula for Tin (IV) Sulfate:** Tin (IV) is Sn^{4+}. Sulfate is SO_4^{2-}. Criss-crossing gives $Sn_2(SO_4)_4$. The subscripts (2 and 4) can be simplified by dividing by 2. The final formula is $\mathbf{Sn(SO_4)_2}$.

5. **Formula for Ammonium Phosphate:** Ammonium is NH_4^+. Phosphate is PO_4^{3-}. Criss-crossing gives $(NH_4)_3PO_4$. The formula is $\mathbf{(NH_4)_3PO_4}$.

Nomenclature Practice (Mixed Types)

Practice: Naming Compounds

1. **Name $\mathbf{Fe_2(SO_4)_3}$:** Ionic. SO_4 is sulfate (-2 charge). Total negative is -6. Two Fe ions must balance this, so each is Fe^{3+}. Name: **Iron (III) sulfate**.

2. **Name $\mathbf{P_2S_5}$:** Molecular. Two phosphorus, five sulfur. Name: **Diphosphorus pentasulfide**.

3. **Name HBr(aq):** Acid. Binary acid (H + nonmetal). Rule: hydro- + (nonmetal root) + -ic acid. Name: **Hydrobromic acid**.

4. **Name $\mathbf{CuSO_4 \cdot {}_5H_2O}$:** Hydrate. Name the ionic part first: $CuSO_4$ is Copper (II) sulfate. The prefix for 5 is "penta-". Name: **Copper (II) sulfate pentahydrate**.

5. **Name $\mathbf{HNO_2}$(aq):** Acid. Oxoacid. The anion is NO_2^- (nitrite). Rule: Anions ending in "-ite" become "-ous acid". Name: **Nitrous acid**.

Practice: Writing Formulas from Names

1. **Formula for Chromium (VI) oxide:** Ionic. Cr^{6+} and O^{2-}. Criss-cross and simplify: $Cr_2O_6 \rightarrow$ $\mathbf{CrO_3}$.

2. **Formula for Dinitrogen trioxide:** Molecular. "di-" = 2 Nitrogen. "tri-" = 3 Oxygen. Formula: $\mathbf{N_2O_3}$.

3. **Formula for Perchloric acid:** Oxoacid. "-ic acid" comes from an "-ate" anion. Perchlorate is ClO_4^-. Balance with H^+. Formula: $\mathbf{HClO_4}$.

4. **Formula for Magnesium sulfate heptahydrate:** Ionic hydrate. Magnesium is Mg^{2+}. Sulfate is SO_4^{2-}. They combine 1:1 to form $MgSO_4$. "hepta-" means 7 waters. Formula: $\mathbf{MgSO_4 \cdot {}_7H_2O}$.

5. **Formula for Vanadium (V) chloride:** Ionic. V^{5+} and Cl^-. Criss-cross. Formula: $\mathbf{VCl_5}$.

Empirical vs. Molecular Formulas

Practice: Deducing Empirical Formulas

1. **Dextrose (Molecular Formula $\mathbf{C_6H_{12}O_6}$):** Divide subscripts (6, 12, 6) by their greatest common divisor (6). Empirical Formula: $\mathbf{CH_2O}$.

2. **Adenine (Molecular Formula $\mathbf{C_5H_5N_5}$):** Divide subscripts (5, 5, 5) by 5. Empirical Formula: **CHN**.

3. **Nitrous Oxide (Molecular Formula $\mathbf{N_2O}$):** Subscripts (2, 1) are already in their simplest whole-number ratio. Empirical Formula: $\mathbf{N_2O}$.

4. **Octane (Molecular Formula $\mathbf{C_8H_{18}}$):** Divide subscripts (8, 18) by their greatest common divisor (2). Empirical Formula: $\mathbf{C_4H_9}$.

5. **Hydrogen Peroxide (Molecular Formula $\mathbf{H_2O_2}$):** Divide subscripts (2, 2) by 2. Empirical Formula: **HO**.

Molar Mass and Percent Composition

- **Molar Mass:** The mass in grams of one mole of a substance (g/mol). Calculated by summing the atomic masses of all atoms in the formula.

- **Percent Composition:** The percentage by mass of each element in a compound.

$$\% \text{ mass of element} = \frac{(n \times \text{molar mass of element})}{\text{molar mass of compound}} \times 100\%$$

Practice: Molar Mass & Percent Composition

1. **Calculate the molar mass of Barium Acetate, $\mathbf{Ba(C_2H_3O_2)_2}$.**

$$1(137.33\,\text{g/mol}) + 4(12.01\,\text{g/mol}) + 6(1.008\,\text{g/mol}) + 4(16.00\,\text{g/mol}) = 255.42\,\text{g/mol}$$

2. **Calculate the mass percent of nitrogen in ammonium nitrate ($\mathbf{NH_4NO_3}$).**
   - **Molar Mass:** $2(14.01\,\text{g/mol}) + 4(1.008\,\text{g/mol}) + 3(16.00\,\text{g/mol}) = 80.05\,\text{g/mol}$
   - **Percent N:** $\frac{2\times 14.01\,\text{g/mol}}{80.05\,\text{g/mol}} \times 100\% = \mathbf{35.0\%}$

3. **Calculate the molar mass of Ibuprofen ($\mathbf{C_{13}H_{18}O_2}$).**

$$13(12.01\,\text{g/mol}) + 18(1.008\,\text{g/mol}) + 2(16.00\,\text{g/mol}) = 206.28\,\text{g/mol}$$

4. **Calculate the mass percent of carbon in propane ($\mathbf{C_3H_8}$).**
   - **Molar Mass:** $3(12.01\,\text{g/mol}) + 8(1.008\,\text{g/mol}) = 44.09\,\text{g/mol}$
   - **Percent C:** $\frac{3\times 12.01\,\text{g/mol}}{44.09\,\text{g/mol}} \times 100\% = \mathbf{81.7\%}$

5. **How many moles are in $25.0\,\text{g}$ of calcium nitrate, $\mathbf{Ca(NO_3)_2}$?**
   - **Molar Mass:** $40.08 + 2(14.01) + 6(16.00) = 164.1\,\text{g/mol}$
   - **Moles:** $25.0\text{ g} \times \frac{1\text{ mol}}{164.1\text{ g}} = \mathbf{0.152\ mol}$

Chapter 8: Chemical Reactions

Balancing Equations and Classifying Reactions

Practice: Balancing and Classifying

1. **_ $\mathbf{Ca(OH)_2(aq)}$ + _ $\mathbf{Al_2(SO_4)_3(aq)}$ -¿ _ $\mathbf{CaSO_4(s)}$ + _ $\mathbf{Al(OH)_3(s)}$**
   - **Balance:** Balance the polyatomic ions first. There are 3 SO_4 on the left, so put a 3 in front of $CaSO_4$. This creates 3 Ca, so put a 3 in front of $Ca(OH)_2$. Now there are 6 OH groups on the left, so put a 2 in front of $Al(OH)_3$ to balance the OH and the Al.
   - **Answer:** $\mathbf{3\,Ca(OH)_2 + Al_2(SO_4)_3 \longrightarrow 3\,CaSO_4 + 2\,Al(OH)_3}$
   - **Classification:** Two ionic compounds swap ions. This is **Double Replacement**.

2. **_ Mg(s) + _ $\mathbf{Fe_2O_3(s)}$ -¿ _ Fe(s) + _ MgO(s)**

- **Balance:** Balance Fe first by placing a 2 in front of Fe. Balance O by placing a 3 in front of MgO. This creates 3 Mg, so place a 3 in front of Mg.
- **Answer:** $\mathbf{3\,Mg(s) + Fe_2O_3(s) \longrightarrow 2\,Fe(s) + 3\,MgO(s)}$
- **Classification:** Magnesium replaces iron in the compound. This is **Single Replacement**.

3. **_ C_2H_4(g) + _ O_2(g) -¿ _ CO_2(g) + _ H_2O(g)**
   - **Balance:** Balance C: 2 CO_2. Balance H: 2 H_2O. Count O on right: $2 \times 2 + 2 \times 1 = 6$. Balance O on left: 3 O_2.
   - **Answer:** $\mathbf{C_2H_4(g) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 2\,H_2O(g)}$
   - **Classification:** Hydrocarbon reacts with oxygen. This is **Combustion**.

4. **_ $PbSO_4$(s) -¿ _ $PbSO_3$(s) + _ O_2(g)**
   - **Balance:** Balance Pb and S (they are 1:1). Balance O. There are 4 on left, 5 on right. Double the lead compounds to get an even number of oxygens.
   - **Answer:** $\mathbf{2\,PbSO_4(s) \longrightarrow 2\,PbSO_3(s) + O_2(g)}$
   - **Classification:** One compound breaks into multiple products. This is **Decomposition**.

5. **_ P_2O_5(s) + _ H_2O(l) -¿ _ H_3PO_4(aq)**
   - **Balance:** Balance P by putting a 2 in front of H_3PO_4. This gives 6 H, so put a 3 in front of H_2O. Check O: 5 + 3 = 8 on left, $2 \times 4 = 8$ on right. It is balanced.
   - **Answer:** $\mathbf{P_2O_5(s) + 3\,H_2O(l) \longrightarrow 2\,H_3PO_4(aq)}$
   - **Classification:** Two reactants form one product. This is **Synthesis (Combination)**.

Stoichiometry and Limiting Reactants

- **Stoichiometry:** Using the mole ratios from a balanced chemical equation to calculate the amounts of reactants consumed or products formed.
- **Limiting Reactant:** The reactant that is completely consumed first in a reaction, thereby limiting the amount of product that can be formed.
- **Theoretical Yield:** The maximum amount of product that can be produced from the given amounts of reactants, calculated based on the limiting reactant.
- **Percent Yield:** The ratio of the actual yield (the amount of product actually obtained) to the theoretical yield, expressed as a percentage.

Practice: Limiting Reactant, Yield, and Excess

1. **Problem:** Given $2\,Na + Br_2 \longrightarrow 2\,NaBr$. If you have 46.0 g of Na and 150.0 g of Br_2, what is the theoretical yield of NaBr?
   - **Molar Masses:** Na = 22.99 g/mol, Br_2 = 159.8 g/mol, NaBr = 102.89 g/mol.
   - **From Na:** $46.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g}} \times \frac{2 \text{ mol NaBr}}{2 \text{ mol Na}} \times \frac{102.89 \text{ g NaBr}}{1 \text{ mol NaBr}} = 206 \text{ g NaBr}$
   - **From Br_2:** $150.0 \text{ g Br}_2 \times \frac{1 \text{ mol Br}_2}{159.8 \text{ g}} \times \frac{2 \text{ mol NaBr}}{1 \text{ mol Br}_2} \times \frac{102.89 \text{ g NaBr}}{1 \text{ mol NaBr}} = 193 \text{ g NaBr}$
   - **Conclusion:** Br_2 is the limiting reactant. The theoretical yield is 193 g **NaBr**.

2. **Using the previous problem, if only 175 g of NaBr were produced, what is the percent yield?**

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \frac{175 \text{ g}}{193 \text{ g}} \times 100\% = \mathbf{90.7\%}$$

3. **Using the same problem, how many grams of the excess reactant (Na) are left over?**
   - **Na Used:** $150.0 \text{ g Br}_2 \times \frac{1 \text{ mol Br}_2}{159.8 \text{ g}} \times \frac{2 \text{ mol Na}}{1 \text{ mol Br}_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 43.2 \text{ g Na used}$

- **Na Left:** 46.0 g initial − 43.2 g used = **2.8 g Na**

4. **For $8\,Ba + S_8 \longrightarrow 8\,BaS$, if you have 50.0 g Ba and 50.0 g S_8, which is the limiting reactant?**
   - **Molar Masses:** Ba = 137.33 g/mol, S_8 = 256.56 g/mol.
   - **Moles Ba:** 50.0 g/137.33 g/mol ≈ 0.364 mol Ba. Moles BaS it can make: 0.364 mol.
   - **Moles S_8:** 50.0 g/256.56 g/mol ≈ 0.195 mol S_8. Moles BaS it can make: $0.195 \times 8 = 1.56$ mol.
   - **Conclusion:** Barium produces fewer moles of product, so **Ba is the limiting reactant**.

5. **For $2\,H_2 + O_2 \longrightarrow 2\,H_2O$, you react 10.0 g H_2 with 64.0 g O_2. What is the theoretical yield of water?**
   - **Molar Masses:** H_2 = 2.016 g/mol, O_2 = 32.00 g/mol, H_2O = 18.02 g/mol.
   - **From H_2:** $10.0\text{ g } H_2 \times \frac{1\text{ mol } H_2}{2.016\text{ g}} \times \frac{2\text{ mol } H_2O}{2\text{ mol } H_2} \times 18.02\frac{\text{g}}{\text{mol}} = 89.4\text{ g } H_2O$
   - **From O_2:** $64.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.00\text{ g}} \times \frac{2\text{ mol } H_2O}{1\text{ mol } O_2} \times 18.02\frac{\text{g}}{\text{mol}} = 72.1\text{ g } H_2O$
   - **Conclusion:** O_2 is limiting. The theoretical yield is 72.1 g **H_2O**.

Chapter 9: Solution Chemistry

Molarity, Dilutions, and pH

Practice Problems: Molarity, Dilution, and pH

1. **Molarity:** A 3.5 mL solution of $Pb(C_2H_3O_2)_4$ contains 0.067 g of solute. What is the molarity?
   - **Molar Mass:** $Pb(C_2H_3O_2)_4$ = 443.4 g/mol.
   - **Moles:** $0.067\text{ g} \times \frac{1\text{ mol}}{443.4\text{ g}} \approx 1.51 \times 10^{-4}$ mol
   - **Volume:** 3.5 mL = 0.0035 L.
   - **Molarity:** $M = \frac{1.51\times10^{-4}\text{ mol}}{0.0035\text{ L}} = \mathbf{0.043\ M}$

2. **Dilution:** How much water must be added to 500.0 mL of a 2.4 M KCl solution to make a 1.0 M solution?
   - **Find Final Volume (V_2):** $V_2 = \frac{M_1V_1}{M_2} = \frac{(2.4\text{ M})(500.0\text{ mL})}{1.0\text{ M}} = 1200\text{ mL}$
   - **Water Added:** This is the difference between the final and initial volumes: 1200 mL−500.0 mL = **700 mL**

3. **pH:** A solution has a pH of 4.120. Find its $[H_3O\equiv]$ concentration.

$$[H_3O\equiv] = 10^{-\text{pH}} = 10^{-4.120} = \mathbf{7.59}\ \times10^{-5}M\mathbf{7.59}\ \times10^{-5}M\mathbf{7.59}\ \times10^{-5}M\mathbf{7.59}\ \times10^{-5}M$$

   (3 decimal places in pH → 3 sig figs in concentration).

4. **Molarity:** How many grams of RbOH (Molar Mass = 102.48 g/mol) are in 35.0 mL of a 5.50 M solution?
   - **Find Moles:** $n = M \times V = (5.50\frac{\text{mol}}{\text{L}}) \times (0.0350\text{ L}) = 0.1925\text{ mol RbOH}$
   - **Find Grams:** $0.1925\text{ mol} \times 102.48\frac{\text{g}}{\text{mol}} = \mathbf{19.7\ g}$

5. **Dilution:** What is the concentration of a solution prepared by diluting 25.0 mL of 18.0 M sulfuric acid to a final volume of 1.50 L?

$$M_2 = \frac{M_1V_1}{V_2} = \frac{(18.0\text{ M})(0.0250\text{ L})}{1.50\text{ L}} = \mathbf{0.300\ M}$$

Topics From Provided Materials NOT on the Exam

Based on the instructions provided in `Exam 2 study topics.txt`, certain topics covered in the Chapter 9 materials are for a future exam and will **not** be on this one. You do not need to study the following:

- **Electrolytes:** The classification of compounds as strong, weak, or nonelectrolytes (`Electrolytes and molecular_ionic equations.pdf`).
- **Solubility Rules:** Memorizing and applying the general solubility rules for ionic compounds in water.
- **Molecular and Ionic Equations:** Writing complete ionic and net ionic equations by identifying spectator ions.
- **Oxidation-Reduction (Redox) Reactions:** All topics from the `Oxidation states and redox reactions.pdf` document, including:
  - Assigning oxidation states (oxidation numbers).
  - Identifying species that are oxidized or reduced.
  - Identifying oxidizing and reducing agents.
  - Balancing redox reactions using the half-reaction method.
  - The Activity Series for predicting single replacement reactions.

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