Comprehensive Study Guide for Chapter 9 Quiz

Quiz Details

- **Date and Time:** Tuesday, October 21st, between 7:00 AM and 7:30 AM.
- **Format:** Closed notes.
- **Allowed Materials:** A scientific calculator.

Key Topics on the Quiz

1. Using the provided solubility rules to determine if an ionic compound is soluble or insoluble.
2. Calculating the pH given [H_3O^+] concentration, applying significant figure rules, and classifying the solution as acidic, basic, or neutral.
3. Stating the oxidation number of an atom in a compound or element (must know the 7 rules by heart).
4. Balancing a redox reaction using either the half-reaction method or the visual method.

Phased Study Plan

Phase 1: Foundation & Memorization (Friday - Saturday)

- **Memorize Oxidation Rules:** Your first priority is to memorize the seven rules for assigning oxidation numbers. This is a non-negotiable requirement for the quiz.
  - **Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slides 48 & 49**. Create flashcards or use mnemonic devices.
- **Understand Core Concepts:** Review the PowerPoint slides to build a strong foundation for each topic.
  - **Solubility:** Understand the terms soluble and insoluble. Review the table on **Slide 34**.
  - **pH Concept:** Understand the logarithmic relationship between pH and [H_3O^+]. Review **Slides 16-19**. Pay special attention to the significant figure rule on Slide 17.
  - **Redox Reactions:** Learn the definitions of oxidation and reduction. Review the half-reaction method on **Slides 58-59**.

Phase 2: Practice & Application (Saturday - Sunday)

- **Solubility Practice:**
  - Work through problems **9.22** and **9.23** on page 404 of the `Questions and Problems` PDF. Use the table on Slide 34 to determine your answers.
- **pH Calculation Drills:**
  - Complete problems **9.80** and **9.81** on page 407 of the `Questions and Problems` PDF. Write down the pH, your reasoning for the number of decimal places, and whether it's acidic or basic.
- **Oxidation Number Drills:**
  - Complete the entire `Oxidation states and redox reactions.pdf` worksheet (Part 1). This is the best way to master the 7 rules.
  - For more practice, solve problems **9.49**, **9.50**, and **9.51** on page 405 of the `Questions and Problems` PDF.

Phase 3: Final Mastery (Monday)

- **Balance Redox Reactions:**
  - Work through all problems in **Part 3** of the `Oxidation states and redox reactions.pdf` worksheet. This gives you direct practice with the half-reaction method.
- **Review Explanations:** Carefully read the detailed walkthroughs in the next section of this guide. Compare your work to the correct methods.
- **Self-Correction:** Redo any problems you got wrong in Phase 2 without looking at the answers. The goal is to be able to solve them correctly and confidently on your own.
- **Final Check:** Get a good night's sleep. Make sure your calculator is ready.

Explanations of Key Topics and Practice Problems

Topic 1: Using Solubility Rules

Goal: To classify an ionic compound as soluble or insoluble based on the provided rules.

- **Reference:** `Chapter 9...pdf`, Slide 34, TABLE 4.1

Example: Classify AgCl and Na_3PO_4.

Step 1: Classify AgCl:

- **Identify Ions:** Ag^+ and Cl^-.
- **Apply Rule:** Find the Cl^- ion in the "Generally Soluble" section. Check its exceptions. The exceptions state that when Cl^- pairs with Ag^+, the compound is insoluble.
- **Conclusion:** AgCl is **insoluble**.

Step 2: Classify Na_3PO_4:

- **Identify Ions:** Na^+ and PO_4^{3-}.
- **Apply Rule:** Find the PO_4^{3-} ion in the "Generally Insoluble" section. Check its exceptions. The exceptions state that when PO_4^{3-} pairs with Na^+ (an alkali metal ion), the compound is soluble.
- **Conclusion:** Na_3PO_4 is **soluble**.

Topic 2: Calculating pH

Goal: To calculate pH from $[H_3O^+]$ with correct significant figures and classify the solution.

- **Reference:** `Chapter 9...pdf`, Slides 17-19

Key Formula and Rule:

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

Sig Fig Rule: The number of **decimal places** in the pH must equal the number of **significant figures** in the $[H_3O^+]$ concentration.

Example Problem (from Slide 19): Calculate the pH for $[H_3O^+] = 8.82 \times 10^{-4}$ M.

Step 1: Identify Significant Figures: The concentration, 8.82×10^{-4}, has **three** significant figures (8, 8, 2).

Step 2: Apply the Rule: The final pH value must be reported to **three** decimal places.

Step 3: Calculate the pH:

$$\text{pH} = -\log(8.82 \times 10^{-4}) = 3.0544...$$

Step 4: Round to Correct Decimal Places: Rounding to three decimal places gives **3.054**.

Step 5: Classify: Since 3.054 is less than 7, the solution is **acidic**.

Topic 3: Stating Oxidation Numbers

Goal: To determine the oxidation number of a specific atom in a compound or ion using the 7 rules.

- **Reference:** `Chapter 9...pdf`, Slides 48-49 (Rules), 52-54 (Examples)

Example Problem (from Slide 53): Determine the oxidation number of Cr in $Na_2Cr_2O_7$.

Step 1: Assign Known Oxidation Numbers:

- Na is a Group 1 metal, so its oxidation number is **+1** (Rule 2 for ions in compounds).
- O has an oxidation number of **-2** (Rule 3).

Step 2: Set Up an Algebraic Equation: The compound is neutral, so the sum of all oxidation numbers must be zero (Rule 6). Let the oxidation number of Cr be 'x'. Since there are two Cr atoms, we use '2x'.

$$\underbrace{2(\mathrm{Na})}_{\text{charge from Na}} + \underbrace{2(\mathrm{Cr})}_{\text{charge from Cr}} + \underbrace{7(\mathrm{O})}_{\text{charge from O}} = 0$$

$$2(+1) + 2(x) + 7(-2) = 0$$

Step 3: Solve for x:

$$\begin{aligned} 2 + 2x - 14 &= 0 \\ 2x - 12 &= 0 \\ 2x &= +12 \\ x &= +6 \end{aligned}$$

Answer: The oxidation number of each chromium (Cr) atom is **+6**.

Topic 4: Balancing Redox Reactions

Goal: To balance a redox equation, ensuring that both atoms and charge are conserved.

- **Reference:** `Chapter 9...pdf`, Slides 58-61

Example Problem (from Slide 59): Balance $\mathbf{Cr(s) + Ni^{2}{\equiv}(aq) \longrightarrow Cr^{3}{\equiv}(aq) + Ni(s)}$

Step 1: Split into Half-Reactions: Identify what is being oxidized (oxidation number increases) and what is being reduced (oxidation number decreases).

- Cr goes from 0 to +3 (loss of e^-), so it is **oxidation**.
- Ni goes from +2 to 0 (gain of e^-), so it is **reduction**.

$$\begin{aligned} \text{Oxidation:} \quad & \mathrm{Cr(s) \longrightarrow Cr^{3}{\equiv}(aq)} \\ \text{Reduction:} \quad & \mathrm{Ni^{2}{\equiv}(aq) \longrightarrow Ni(s)} \end{aligned}$$

Step 2: Balance Charge with Electrons: Add electrons to the more positive side of each equation.

$$\begin{aligned} \text{Oxidation:} \quad & \mathrm{Cr(s) \longrightarrow Cr^{3}{\equiv}(aq) + {}_{3}e^{-}} \\ \text{Reduction:} \quad & \mathrm{Ni^{2}{\equiv}(aq) + {}_{2}e^{-} \longrightarrow Ni(s)} \end{aligned}$$

Step 3: Equalize Electrons: The least common multiple of 3 electrons (lost) and 2 electrons (gained) is 6.

- Multiply the oxidation half-reaction by **2**.
- Multiply the reduction half-reaction by **3**.

$$\begin{aligned} \text{Oxidation } (\times 2): \quad & \mathrm{{}_{2}Cr(s) \longrightarrow {}_{2}Cr^{3}{\equiv}(aq) + {}_{6}e^{-}} \\ \text{Reduction } (\times 3): \quad & \mathrm{{}_{3}Ni^{2}{\equiv}(aq) + {}_{6}e^{-} \longrightarrow {}_{3}Ni(s)} \end{aligned}$$

Step 4: Combine and Cancel: Add the two new half-reactions together and cancel the species that appear on both sides (the electrons).

$$_2Cr(s) + {}_3Ni^2{\equiv}(aq) + {}_6e^- \longrightarrow {}_2Cr^3{\equiv}(aq) + {}_3Ni(s) + {}_6e^-$$

Final Balanced Equation:

$$_2Cr(s) + {}_3Ni^2{\equiv}(aq) \longrightarrow {}_2Cr^3{\equiv}(aq) + {}_3Ni(s)$$

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