Study Guide for Chemistry Quiz (Chapter 9 Topics)

Quiz Date: Tuesday, October 21st

This guide is tailored specifically for the four topics on your upcoming quiz. It includes key concepts, rules to memorize, and worked examples with references to your course materials.

Topic 1: Using Solubility Rules

Key Concept

You must be able to use the provided solubility rules to determine if an ionic compound will dissolve in water (**soluble**) or form a solid precipitate (**insoluble**).

- **Primary Study Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slide 34 (TABLE 4.1)**.
- **Soluble compounds** dissociate completely into ions in water. They are strong electrolytes.
- **Insoluble compounds** do not dissolve in water. They are considered weak electrolytes because a negligible amount of ions are formed.

How to Apply the Rules

1. **Identify the ions** that make up the compound (one cation, one anion).
2. **Check the "Generally Soluble" Section First:** Find one of the ions in this section.
   - If the other ion in your compound is **NOT** listed as an exception, the compound is **SOLUBLE**.
   - If the other ion **IS** listed as an exception, the compound is **INSOLUBLE**.
3. **Check the "Generally Insoluble" Section if Needed:** If your ion was not in the first section, find it here.
   - If the other ion in your compound is **NOT** listed as an exception, the compound is **INSOLUBLE**.
   - If the other ion **IS** listed as an exception, the compound is **SOLUBLE**.

Practice Problems

- **Practice Source:** `Questions and Problems`, Page 404, Problems **9.22 & 9.23**.
- **Practice Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, Slides **36 & 37**.

1. **Is $CaCO_3$ soluble or insoluble?**
   - **Ions:** Ca^{2+} and CO_3^{2-}.
   - **Rule:** The CO_3^{2-} ion is in the "Generally Insoluble" category. Ca^{2+} is not listed as an exception.
   - **Answer:** Therefore, $CaCO_3$ is **insoluble**.
2. **Is K_2S soluble or insoluble?**

- **Ions:** K^+ and S^{2-}.
- **Rule:** The S^{2-} ion is "Generally Insoluble". However, K^+ is an alkali metal ion and is listed as an exception.
- **Answer:** Therefore, K_2S is **soluble**.

3. **Is $PbCl_2$ soluble or insoluble?**

   - **Ions:** Pb^{2+} and Cl^-.
   - **Rule:** The Cl^- ion is "Generally Soluble". However, Pb^{2+} is listed as an exception.
   - **Answer:** Therefore, $PbCl_2$ is **insoluble**.

Topic 2: pH Calculations

Key Concepts and Formulas

You must be able to calculate pH from $[H_3O^+]$ and determine if the solution is acidic, basic, or neutral, while applying the correct significant figure rules.

- **Primary Study Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slides 16-19**.
- **Formula to find pH:**
$$\mathrm{pH} = -\log[\mathrm{H_3O^+}]$$
- **pH Scale:** At 25 °C, a pH < 7 is **acidic**, pH = 7 is **neutral**, and pH > 7 is **basic**.

Crucial Rule for Significant Figures

- The number of **significant figures** in your $[H_3O^+]$ concentration value determines the number of **decimal places** in your final pH answer.

Practice Problems

- **Practice Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slide 19**.
- **Practice Source:** `Questions and Problems`, Page 407, Problems **9.80 & 9.81**.

1. **Calculate the pH for $[H_3O^+]$ =** 4.3×10^{-8} M **and classify the solution.**
   - **Calculation:** $\mathrm{pH} = -\log(4.3 \times 10^{-8})$
   - **Sig Figs:** The concentration 4.3×10^{-8} has **two** significant figures. Therefore, the pH value must have **two** decimal places.
   - **Answer:** pH = **7.37**. Since 7.37 ¿ 7, the solution is **basic**.
2. **Calculate the pH for $[H_3O^+]$ =** 8.82×10^{-4} M **and classify the solution.**
   - **Calculation:** $\mathrm{pH} = -\log(8.82 \times 10^{-4})$
   - **Sig Figs:** The concentration 8.82×10^{-4} has **three** significant figures. Therefore, the pH value must have **three** decimal places.
   - **Answer:** pH = **3.054**. Since 3.054 ¡ 7, the solution is **acidic**.

Topic 3: Assigning Oxidation Numbers

Key Concept

You must memorize and apply the seven rules for assigning oxidation numbers to an atom in a compound or ion.

The Seven Rules (Memorize These)

- **Primary Study Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slides 48 & 49**.

Rule 1: Elements: The oxidation number of an atom in its elemental form is **0** (e.g., Na, O_2, S_8).

Rule 2: Monatomic Ions: The oxidation number equals the **charge** of the ion (e.g., Na^+ is +1; Cl^- is -1).

Rule 3: Oxygen: The oxidation number is usually **-2**. (Exception: -1 in peroxides like H_2O_2).

Rule 4: **Hydrogen:** The oxidation number is **+1** with nonmetals and **-1** with metals.

Rule 5: **Halogens:** The oxidation number is usually **-1**. (Exception: positive when bonded to oxygen or a more electronegative halogen). Fluorine is **always -1**.

Rule 6: **Neutral Compounds:** The **sum** of all oxidation numbers must equal **zero**.

Rule 7: **Polyatomic Ions:** The **sum** of all oxidation numbers must equal the **charge of the ion**.

Practice Problems

- **Practice Source:** `Oxidation states and redox reactions.pdf` (worksheet).
- **Practice Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slides 50-54**.
- **Practice Source:** `Questions and Problems`, Page 405, Problems **9.49 - 9.51**.

1. **Determine the oxidation number for Mn in $KMnO_4$.**
   - **Rules:** K is +1 (Group 1 ion). O is -2 (Rule 3). The sum must be 0 (Rule 6). Let Mn = x.
   - **Equation:** $(+1) + (x) + 4(-2) = 0$
   - **Solve:** $1 + x - 8 = 0 \implies x = +7$. The oxidation number of Mn is **+7**.

2. **Determine the oxidation number for N in NO_3^-.**
   - **Rules:** O is -2 (Rule 3). The sum must be -1 (Rule 7). Let N = x.
   - **Equation:** $(x) + 3(-2) = -1$
   - **Solve:** $x - 6 = -1 \implies x = +5$. The oxidation number of N is **+5**.

Topic 4: Balancing Redox Reactions

Key Concepts

A redox reaction involves the transfer of electrons, identified by changes in oxidation numbers. You must be able to balance the atoms and the charge.

- **Primary Study Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slides 58-61**.
- **Oxidation:** Loss of electrons (oxidation number increases).
- **Reduction:** Gain of electrons (oxidation number decreases).

The Half-Reaction Method

Step 1: Split into Half-Reactions: Write separate equations for the oxidation and reduction processes.

Step 2: Balance Atoms: Balance all atoms other than O and H.

Step 3: Balance Charge with Electrons: Add electrons (e^-) to the more positive side of each equation to make the charges equal on both sides.

Step 4: Equalize Electrons: Multiply the half-reactions by integers to make the number of electrons lost in oxidation equal the number of electrons gained in reduction.

Step 5: Combine: Add the balanced half-reactions and cancel the electrons.

Practice Problem

- **Practice Source:** `Oxidation states and redox reactions.pdf`, **Part 3**.
- **Practice Source:** `Chapter 9 Chemical reactions in aqueous solutions.pdf`, **Slides 59 & 61.

Balance the following redox reaction: $\mathbf{Sn(s) + H\equiv(aq) \longrightarrow Sn^{2}\equiv(aq) + H_2(g)}$

- **Step 1: Split into Half-Reactions**

$$\begin{aligned} \text{Oxidation:} \quad & \mathrm{Sn(s) \longrightarrow Sn^{2}\equiv(aq)} \\ \text{Reduction:} \quad & \mathrm{H\equiv(aq) \longrightarrow H_2(g)} \end{aligned}$$

- **Step 2: Balance Atoms**

$$\begin{aligned} \text{Oxidation:} \quad & \mathrm{Sn(s) \longrightarrow Sn^{2}\equiv(aq)} \quad \text{(Sn is balanced)} \\ \text{Reduction:} \quad & \mathrm{{}_2H\equiv(aq) \longrightarrow H_2(g)} \quad \text{(H is now balanced)} \end{aligned}$$

- **Step 3: Balance Charge with Electrons**

$$\begin{aligned} \text{Oxidation:} \quad & \mathrm{Sn(s) \longrightarrow Sn^{2}\equiv(aq) + {}_2e^-} \quad \text{(Charge is 0 on both sides)} \\ \text{Reduction:} \quad & \mathrm{{}_2H\equiv(aq) + {}_2e^- \longrightarrow H_2(g)} \quad \text{(Charge is 0 on both sides)} \end{aligned}$$

- **Step 4: Equalize Electrons**
  - The oxidation half-reaction loses 2 electrons.
  - The reduction half-reaction gains 2 electrons.
  - The electrons are already balanced (2 = 2). No multiplication is needed.

- **Step 5: Combine**
  - Add the two equations:

$$\mathrm{Sn(s)} + {}_2\mathrm{H}{\equiv}\mathrm{(aq)} + {}_2\mathrm{e}^- \longrightarrow \mathrm{Sn}^2{\equiv}\mathrm{(aq)} + {}_2\mathrm{e}^- + \mathrm{H_2(g)}$$

  - Cancel the electrons (${}_2\mathrm{e}^-$) from both sides.
  - **Final Balanced Equation:**

$$\mathrm{Sn(s)} + {}_2\mathrm{H}{\equiv}\mathrm{(aq)} \longrightarrow \mathrm{Sn}^2{\equiv}\mathrm{(aq)} + \mathrm{H_2(g)}$$