

Redox Stoichiometry (Titrations)

- In a redox titration, a known concentration of an oxidizing agent is used to find an unknown concentration of a reducing agent, or vice-versa.
- Applications include finding the iron content of water or the Vitamin C content in foods.

General Steps:

1. Balance the redox equation
2. Organize data for the given and required substances
3. Find the moles of the given substance
If given volume and concentration, use $n = Cv$
If given mass, find molar mass, then use $n = m/M$
4. Find the moles of the required substances, using mole ratio
5. Convert moles to concentration ($C = n/v$) or volume ($v = n/C$) or mass ($m = nM$) of required substance

Example 1:

The following example involves the titration of a solution of sodium oxalate, $\text{Na}_2\text{C}_2\text{O}_4(\text{aq})$ with a solution of potassium permanganate, $\text{KMnO}_4(\text{aq})$.

Important Note:

- The permanganate ion, MnO_4^- , is often used as an oxidizing agent because of its deep purple color.
- It acts as its own indicator.
- It is reduced to pale pink Mn^{2+} by the oxalate ions.
- Once all oxalate ions have been oxidized the next drop of permanganate ions turns the solution a faint purple color, signaling the endpoint of the titration.

Find the **volume** of 0.200 M solution of MnO_4^- which will react with 50.0 mL of 0.100 M solution of $\text{C}_2\text{O}_4^{2-}$.



Example 2:

A 23.30 ml sample of KMnO_4 solution is decolorized by 0.1111g of oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$). The products are Mn^{2+} and CO_2 gas. Calculate the concentration of KMnO_4 using the following equation:

Example 3:

The oxidation of Fe^{2+} ions by permanganate in acidic solution is accompanied by the formation of Fe^{3+} ions and Mn^{2+} ions. How many moles of FeSO_4 would be oxidized by 100.0 ml of 0.02118 mol/L KMnO_4 solution.



Redox Stoichiometry Worksheet

R Bogdan

A red circular graphic with a gradient, appearing as a semi-circle or a partial circle, located to the right of the author's name.

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Table of Contents Redox Stoichiometry Worksheet

1. Understanding the eBook Redox Stoichiometry Worksheet
 - The Rise of Digital Reading Redox Stoichiometry Worksheet
 - Advantages of eBooks Over Traditional Books
2. Identifying Redox Stoichiometry Worksheet
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Redox Stoichiometry Worksheet
 - User-Friendly Interface
4. Exploring eBook Recommendations from Redox Stoichiometry Worksheet
 - Personalized Recommendations
 - Redox Stoichiometry Worksheet User Reviews and Ratings
 - Redox Stoichiometry Worksheet and Bestseller Lists
5. Accessing Redox Stoichiometry Worksheet Free and Paid eBooks
 - Redox Stoichiometry Worksheet Public Domain eBooks
 - Redox Stoichiometry Worksheet eBook Subscription Services
 - Redox Stoichiometry Worksheet Budget-Friendly Options

6. Navigating Redox Stoichiometry Worksheet eBook Formats
 - ePub, PDF, MOBI, and More
 - Redox Stoichiometry Worksheet Compatibility with Devices
 - Redox Stoichiometry Worksheet Enhanced eBook Features
7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Redox Stoichiometry Worksheet
 - Highlighting and Note-Taking Redox Stoichiometry Worksheet
 - Interactive Elements Redox Stoichiometry Worksheet
8. Staying Engaged with Redox Stoichiometry Worksheet
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Redox Stoichiometry Worksheet
9. Balancing eBooks and Physical Books Redox Stoichiometry Worksheet
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Redox Stoichiometry Worksheet
10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Redox Stoichiometry Worksheet
 - Setting Reading Goals Redox Stoichiometry Worksheet
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Redox Stoichiometry Worksheet
 - Fact-Checking eBook Content of Redox Stoichiometry Worksheet
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
14. Embracing eBook Trends
 - Integration of Multimedia Elements

- Interactive and Gamified eBooks

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