What Is The Coefficient Of The Permanganate Ion When The Following Equation Is Balanced?MnO4- + Br- ?

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Balancing chemical equations is a fundamental skill in chemistry that allows scientists to understand the stoichiometry of reactions accurately. Specifically, when dealing with redox (reduction-oxidation) reactions, such as the interaction between permanganate ions (MnO_4^-) and bromide ions (Br^-), it is crucial to determine the correct coefficients that balance both mass and charge. This process involves understanding oxidation states, electron transfer, and applying the principles of conservation of mass and charge. In this article, we will explore the step-by-step method to find the coefficients in the balanced equation, focusing on the coefficients of MnO_4^-, Br^-, and other reactants and products involved.

Understanding the Reactants and Their Roles in the Reaction

Permanganate Ion (MnO_4^-)

The permanganate ion is a well-known oxidizing agent in aqueous solutions. It contains manganese in a high oxidation state of +7. When it reacts with a reducing agent like bromide ions, manganese is typically reduced to a lower oxidation state, often +2 in acidic conditions.

Bromide Ion (Br^-)

Bromide ions are usually oxidized during reactions involving permanganate. In acidic solutions, bromide is oxidized to bromine (Br_2), which is a diatomic molecule.

Determining the Oxidation States

Before balancing the equation, identify oxidation states:

- Mn in MnO_4^-: +7
- Br in Br^-: -1
- Br in Br_2: 0 (elemental form)

This indicates that bromide ions are oxidized from -1 to 0, and permanganate ions are reduced from +7 to +2 (commonly in acidic solutions).

Writing the Unbalanced Ionic Equation

The unbalanced reaction in acidic solution can be written as:

$MnO_4^- + Br^- \rightarrow Mn^{2+} + Br_2$

Note that in an acidic medium, Mn^{2+} and Br_2 are typical products after reduction and oxidation, respectively.

Balancing the Redox Equation Step-by-Step

Step 1: Separate the Oxidation and Reduction Half-Reactions

Identify the two half-reactions:

Oxidation (Bromide to Bromine):

$Br^- \rightarrow Br_2$

Reduction (Permanganate to Manganese ion):

$MnO_4^- \rightarrow Mn^{2+}$

Step 2: Balance Atoms Other Than O and H

For the oxidation half-reaction:

- Balance bromine atoms:

$2\ Br^- \rightarrow Br_2$

For the reduction half-reaction:

- Manganese is already balanced:

$MnO_4^- \rightarrow Mn^{2+}$

Step 3: Balance Oxygen Atoms Using H_2O

In the reduction half-reaction:

MnO_4^- contains 4 oxygen atoms. To balance oxygen:

$MnO_4^- \rightarrow Mn^{2+} + 4\ H_2O$ (on the product side)

But since Mn^{2+} is in the product, and oxygen is in MnO_4^-, the balanced reduction half-reaction is:

$MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

This step introduces H^+ ions to balance oxygen in acidic solution.

Step 4: Balance Hydrogen Atoms Using H^+

Hydrogen atoms are balanced by adding H^+ ions to the side deficient in H:

$MnO^-_4 + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

Step 5: Balance Charge Using Electrons

Determine the change in charge:

- Left side charge: MnO^-_4 (-1) + 8 H^+ (+8) + 5 e^- (-5) = (+2)
- Right side charge: Mn^{2+} (+2)

Electrons are added to balance charge; here, 5 electrons are needed for reduction.

For oxidation, balance the half-reaction:

$2\ Br^- \rightarrow Br_2 + 2\ e^-$

Since bromide is oxidized from -1 to 0, electrons are produced.

Step 6: Equalize Electrons in Both Half-Reactions

Multiply the oxidation half-reaction by 5 and the reduction half-reaction by 2 to have the same number of electrons:

Oxidation:
$10\ Br^- \rightarrow 5\ Br_2 + 10\ e^-$

Reduction:
$2\ MnO^-_4 + 16\ H^+ + 10\ e^- \rightarrow 2\ Mn^{2+} + 8\ H_2O$

Step 7: Add the Half-Reactions Together

Combine:

$$2\ MnO_4^- + 16\ H^+ + 10\ e^- + 10\ Br^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Br_2 + 10\ e^-$$

Electrons cancel out:

$$2\ MnO_4^- + 16\ H^+ + 10\ Br^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Br_2$$

The Balanced Overall Equation and Coefficients

The final balanced reaction in acidic solution is:

$$2\ MnO_4^- + 16\ H^+ + 10\ Br^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Br_2$$

From this, the coefficients are:

- Permanganate ion (MnO_4^-): 2
- Bromide ion (Br^-): 10
- Hydrogen ion (H^+): 16
- Manganese ion (Mn^{2+}): 2
- Bromine (Br_2): 5
- Water (H_2O): 8

Answer to the question: The coefficient of the permanganate ion, MnO_4^-, when the equation is balanced, is 2.

Additional Considerations and Variations

Balancing in Neutral or Basic Solutions

While the above process is specific to acidic conditions, similar principles apply to neutral or basic solutions, with adjustments such as adding OH^- ions instead of H^+ to balance hydrogen and oxygen atoms.

Implications in Real-World Chemistry

Understanding these coefficients is essential in titrations, environmental chemistry, and industrial processes where permanganate and bromide are involved, such as water treatment and analytical chemistry.

Summary

- The key steps involve separating into half-reactions, balancing atoms and charges, and combining the equations.
- The balanced equation reveals the coefficients, with 2 for MnO_4^-.
- Proper understanding of oxidation states and electron transfer ensures accurate balancing in redox reactions.

Conclusion:
The coefficient of the permanganate ion (MnO_4^-) in the balanced redox equation with bromide ions (Br^-) under acidic conditions is 2. Mastery of balancing techniques allows chemists to accurately predict product quantities, analyze reaction mechanisms, and apply this knowledge across various scientific disciplines.

Frequently Asked Questions

What is the coefficient of permanganate ion (MnO_4^-) when balancing the reaction with bromide ion (Br^-) in acidic solution?

The coefficient of MnO4- is 1 when the reaction between permanganate and bromide ions is balanced in acidic solution.

How do you determine the coefficient of MnO_4^- in the redox reaction with Br^-?

By balancing the redox reaction in acidic solution, you assign oxidation states, write half-reactions, balance atoms and charge, and then combine them to find the coefficients, which typically results in a coefficient of 1 for MnO4-.

What is the balanced chemical equation for MnO_4^- reacting with Br^- in acidic medium?

The balanced equation is: MnO4- + 8 Br- + 8 H+ □ Mn2+ + 4 Br2 + 4 H2O, where the coefficient of MnO4- is 1.

Why is the coefficient of MnO_4^- equal to 1 in the balanced reaction with bromide ions?

Because the balancing process shows that one permanganate ion reacts with eight bromide ions, resulting in a coefficient of 1 for MnO4-.

Does the coefficient of MnO_4^- change if the reaction occurs in neutral

or basic solution?

Yes, the coefficients can change depending on the medium; in neutral or basic solutions, the balanced equations differ, but in acidic solution, the coefficient of MnO4- remains 1.

Is the coefficient of MnO4- always 1 when reacting with halide ions like Br-?

Not necessarily; it depends on the stoichiometry of the balanced redox reaction. In the case of MnO4- reacting with Br-, it is 1 in the balanced acidic solution.

What role does the oxidation state of manganese play in determining the coefficients in the reaction?

The change in oxidation state of manganese from +7 in MnO4- to +2 in Mn2+ helps determine the number of electrons transferred and thus influences the coefficients during balancing.

Can you explain the steps to find the coefficient of MnO4- in this redox reaction?

Yes, first write the unbalanced reaction, determine oxidation states, write half-reactions, balance atoms and charge in each half-reaction, then combine and balance electrons to find the coefficients, which give MnO4- a coefficient of 1 in the final balanced equation.

Additional Resources

Coefficient of the Permanganate Ion in the Balanced Equation

Understanding the stoichiometry of redox reactions is fundamental in chemistry, especially when dealing with transition metal ions like permanganate (MnO_4^-). Determining the coefficient of the permanganate ion in a balanced chemical equation involving bromide ions (Br^-) not only elucidates

the reaction mechanism but also reinforces core concepts of oxidation-reduction processes. This article delves deeply into the principles, procedures, and step-by-step calculations necessary to identify the coefficient of MnO_4^-, providing clarity for students, educators, and chemistry enthusiasts alike.

Introduction to Redox Reactions and the Role of Permanganate Ion

Redox (reduction-oxidation) reactions are fundamental to chemistry, underpinning processes ranging from biological metabolism to industrial synthesis. Permanganate ion (MnO_4^-) is a well-known oxidizing agent, especially in aqueous solutions, due to its high oxidation state of manganese (Mn). It is frequently used in titrations and analytical chemistry because of its vivid purple color and strong oxidizing properties.

Permanganate as an Oxidizing Agent:

- Oxidation State of Mn in MnO_4^-: +7
- Color in Solution: Deep purple
- Common Uses: Oxidizing agents in titrations, water treatment, organic synthesis

In reactions with bromide ions (Br^-), MnO_4^- can oxidize Br^- to bromine (Br_2), while itself being reduced from Mn(+7) to a lower oxidation state, typically Mn(II).

Understanding the Reaction: MnO_4^- and Br^-

The fundamental reaction between permanganate ions and bromide ions in acidic solution is a classic example of redox chemistry. The general unbalanced reaction can be written as:

Unbalanced Equation:

$MnO_4^- + Br^- \rightarrow Mn^{2+} + Br_2$

However, to determine the coefficients, especially for MnO_4^-, a systematic approach involving oxidation states, half-reactions, and balancing techniques is essential.

Step-by-Step Procedure for Balancing the Equation

1. Identify Oxidation States

- Manganese in MnO_4^-: +7 (since oxygen is usually -2, and the overall charge is -1)
- Bromide ion (Br^-): -1
- Manganese in Mn^{2+}: +2
- Bromine in Br_2: 0 (elemental form)

This indicates:

- Mn is reduced from +7 to +2.
- Br is oxidized from -1 to 0.

2. Write the Half-Reactions

Oxidation Half-Reaction (Bromide to Bromine):

$Br^- \rightarrow Br_2$

Reduction Half-Reaction (Permanganate to Mn^{2+}):

$MnO_4^- \rightarrow Mn^{2+}$

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3. Balance Each Half-Reaction Separately

Balancing the Oxidation Half-Reaction:

- First, balance Br atoms:

$2\ Br^- \rightarrow Br_2$

- Balance charge by adding electrons:

$2\ Br^- \rightarrow Br_2 + 2\ e^-$

Balancing the Reduction Half-Reaction:

- Balance Mn:

$MnO_4^- \rightarrow Mn^{2+}$

- Balance oxygen by adding H_2O:

$MnO_4^- \rightarrow Mn^{2+} + 4\ H_2O$

- Balance hydrogen by adding H^+ (assuming acidic solution):

Add 8 H^+ to the left:

$MnO_4^- + 8\ H^+ \rightarrow Mn^{2+} + 4\ H_2O$

- Balance charge by adding electrons:

Calculate total charge:

Left: (-1) + (+8) = +7

Right: +2

Electrons needed: 5 electrons to balance the charge:

$MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

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4. Combine the Half-Reactions

- The electrons in oxidation and reduction half-reactions must be equal:

Oxidation: $2\ Br^- \rightarrow Br_2 + 2\ e^-$

Reduction: $MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

- To equalize electrons, multiply the oxidation half-reaction by 5 and the reduction half-reaction by 2:

Oxidation:

$(5)(2\ Br^- \rightarrow Br_2 + 2\ e^-) \rightarrow 10\ Br^- \rightarrow 5\ Br_2 + 10\ e^-$

Reduction:

$(2)(MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O) \rightarrow 2\ MnO_4^- + 16\ H^+ + 10\ e^- \rightarrow 2\ Mn^{2+} + 8\ H_2O$

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5. Write the Final Balanced Equation

- Combine the two:

$2\ MnO_4^- + 16\ H^+ + 10\ Br^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Br_2$

- The electrons cancel out, confirming the balancing.

Determining the Coefficient of MnO_4^-

The key insight from the balanced equation is the coefficient in front of MnO_4^-.

Balanced Equation:

$2\ MnO_4^- + 16\ H^+ + 10\ Br^- \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ Br_2$

Thus, the coefficient of MnO_4^- is 2.

Implications and Applications

Why is this important?

- Quantitative Analysis: Knowing the coefficients allows chemists to determine molar ratios, useful in titration calculations and stoichiometric conversions.
- Reaction Mechanisms: Understanding the transfer of electrons and the balanced equations provides insight into reaction pathways.
- Industrial Processes: Such reactions underpin processes like water purification, organic synthesis, and analytical titrations involving permanganate.

In practice, if you are asked to find the coefficient of MnO_4^- in the reaction with Br_2, you now have a clear, step-by-step methodology rooted in fundamental principles.

Summary

- The balanced redox reaction of permanganate ion (MnO_4^-) with bromide ions (Br_2) in acidic solution involves multiple steps: identifying oxidation states, writing half-reactions, balancing electrons, and combining reactions.
- The coefficient of MnO_4^- in the balanced equation is 2.

- This ratio reflects the stoichiometry of the reaction, ensuring mass and charge balance.

Final Remarks

Mastering the balancing of redox equations like MnO_4^- + Br^- is more than an academic exercise; it is a vital skill in chemical analysis and industrial chemistry. The systematic approach outlined here offers a robust framework for tackling similar problems involving other oxidizing agents and reducing agents. Whether in laboratory titrations, environmental chemistry, or research and development, understanding the coefficients in such reactions is essential for precise calculations and meaningful scientific insights.

In conclusion, the coefficient of the permanganate ion (MnO_4^-) when the equation is balanced with bromide ions in acidic solution is 2. This value emerges from meticulous half-reaction balancing and electron transfer considerations, encapsulating core principles of redox chemistry and stoichiometry.

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