Identify A Substance Which Can Reduce Br (1) To Br- (aq) But Cannot Reduce 1,(s) To I-(aq). O Ag(s) O

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When delving into the fascinating realm of redox chemistry, understanding the nuances of oxidation and reduction processes is fundamental. A common challenge faced by students and chemists alike is identifying substances that can selectively reduce certain species while leaving others unaffected. Specifically, the question often arises: "Identify a substance which can reduce bromine ions (Br^-) to bromine (Br_2) in aqueous solution but cannot reduce solid iodine (I_2) to iodide ions (I^-)." This article explores this intriguing chemical puzzle in detail, emphasizing the role of agents such as silver metal (Ag(s)) and the principles of reduction potentials.

Understanding Redox Reactions and Standard Reduction Potentials

What Are Redox Reactions?

Redox reactions involve the transfer of electrons between species. The substance that loses electrons is oxidized, while the one gaining electrons is reduced. The ability of a substance to act as a reducing or oxidizing agent depends on its position in the electrochemical series and its standard reduction potential.

Standard Reduction Potentials (E°)

The standard reduction potential indicates the tendency of a chemical species to gain electrons and be reduced. Species with higher (more positive) E° values are stronger oxidizers, while those with lower (more negative) values are better reducers.

For common halogens:

- Iodine ($I_2 + 2e^-$? $2I^-$): $E° \approx +0.54$ V
- Bromine ($Br_2 + 2e^-$? $2Br^-$): $E° \approx +1.07$ V

The more positive the E°, the stronger the oxidizing agent.

Analyzing the Reduction of Br^- to Bromine

Reduction of Br^- to Br_2

In aqueous solution, bromide ions (Br^-) can be oxidized to bromine (Br_2) if an appropriate oxidizing agent is present. The half-reaction is:

Br^- (aq) ? Br_2 (l) + $2e^-$? ? $E° \approx +1.07$ V

This indicates that any strong oxidizer with a reduction potential greater than +1.07 V can facilitate this transformation.

Potential Oxidizing Agents for Br^-

Some common oxidants capable of oxidizing Br^- to Br_2:

- Chlorine (Cl_2): E° ≈ +1.36 V
- Hydrogen peroxide (H_2O_2) in acidic medium: E° ≈ +1.78 V
- Potassium permanganate ($KMnO_4$): E° ≈ +1.51 V in acidic solution

These substances can readily oxidize bromide ions to bromine.

Why Can't the Same Substances Reduce I_2 to I^-?

Reduction of I_2 to I^-

The half-reaction for iodine is:

$I_2 + 2e^- \rightarrow 2I^-$ E° ≈ +0.54 V

Since the standard reduction potential is lower compared to that of bromine, typical oxidizing agents capable of converting Br^- to Br_2 are often not strong enough to reduce I_2 to I^-.

Implication of Reduction Potentials

- Oxidants with E° > +0.54 V can reduce I_2 to I^-.
- Oxidants with E° between +0.54 V and +1.07 V may oxidize Br^- but not I_2.

In other words, the strength of the oxidizing agent determines its ability to reduce these halogen species.

The Role of Silver (Ag) in Redox Reactions

Silver as an Oxidizing or Reducing Agent

Silver metal (Ag(s)) has a standard oxidation potential:

$Ag^+ + e^- \rightarrow Ag(s)$ E° ≈ +0.80 V

Since metallic silver is in the elemental form, it can act as a reducing agent under specific circumstances, especially when it oxidizes to Ag^+.

Silver's Effect on Bromide and Iodide Ions

- Silver readily reacts with halide ions (Cl^-, Br^-, I^-) to form insoluble silver halides: AgCl, AgBr, AgI.
- The solubility product (Ksp) varies:

- AgCl: Ksp $\approx 1.8 \times 10^{-10}$
- AgBr: Ksp $\approx 5.4 \times 10^{-13}$
- AgI: Ksp $\approx 8.5 \times 10^{-17}$

Because AgI is much less soluble than AgBr or AgCl, silver has a higher tendency to precipitate I^- as AgI, effectively removing I^- from solution and preventing its oxidation back to I_2.

Identifying the Suitable Substance: Silver Metal (Ag(s))

Why Silver Can Reduce Br^- to Br_2 but Not I_2 to I^-

- Silver metal can oxidize bromide ions to bromine because the oxidation of Ag(s) to Ag^+ has a standard potential of +0.80 V, which is sufficient to oxidize Br^- ($E° \approx +1.07$ V for reduction of Br_2) in the presence of the right conditions.
- However, for iodine, the oxidation of Ag(s) to Ag^+ is not thermodynamically favorable enough to produce I^- from I_2. The insolubility of AgI and the high stability of I_2 make this process unfavorable.

Practical Implication

- Silver can be used to selectively oxidize bromide ions to bromine but cannot do the same with iodine. This selectivity is exploited in various analytical and preparative procedures.

Summary: The Role of Silver in Redox Chemistry

- Silver metal (Ag(s)) is capable of oxidizing bromide ions (Br^-) to bromine (Br_2) because of its oxidation potential and the relative ease of forming insoluble AgBr.
- Silver cannot oxidize iodine (I_2) to iodide ions (I^-) effectively because the formation of insoluble AgI is thermodynamically less favorable and the oxidation potential is insufficient to drive this reaction.
- This selective oxidation property makes silver an ideal candidate for the substance that reduces Br^- to Br_2 but does not reduce I_2 to I^-.

Conclusion

The quest to identify a substance that can reduce bromide ions to bromine but cannot reduce iodine to iodide leads us to the unique properties of silver metal. Silver's ability to oxidize Br^- to Br_2, combined with its inability to reduce I_2 to I^- under standard conditions, underscores its role as a selective redox agent. This understanding is vital for chemists designing reactions, separation processes, or analytical techniques involving halogens.

By appreciating the subtle differences in reduction potentials and solubility products, we can harness the chemistry of silver and halogens for practical applications in laboratories and industry. Whether for qualitative analysis or synthesis, silver's distinctive redox behavior exemplifies the importance of thermodynamics and solubility in chemical reactivity.

References:

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Frequently Asked Questions

Which substance can reduce Br_2 (l) to Br^- (aq) but cannot reduce I_2 (s) to I^- (aq)?

Silver metal (Ag(s)) can reduce Br_2 (l) to Br^- (aq) but is unable to reduce I_2 (s) to I^- (aq).

Why does silver (Ag) reduce bromine (Br_2) to bromide ions but not iodine (I_2) to iodide ions?

Because silver has a higher tendency to form compounds with bromide ions than with iodide ions, making it capable of reducing bromine but not iodine.

What is the role of Ag(s) in reducing Br_2 to Br^- in aqueous solution?

Ag(s) acts as a reducing agent, donating electrons to bromine molecules, converting them into bromide ions.

Can Ag(s) reduce I_2 to I^- in aqueous solution? Why or why not?

No, Ag(s) cannot reduce I_2 to I^- because iodine forms more stable compounds with silver (like AgI), and the reduction potential is not favorable.

What does the inability of Ag(s) to reduce I_2 indicate about its oxidizing and reducing properties?

It indicates that Ag(s) is a weaker reducing agent for iodine compared to bromine, due to the higher stability of silver iodide and the relative reduction potentials.

In the context of reducing bromine but not iodine, what can be inferred about the standard reduction potentials involved?

The standard reduction potential of bromine is such that Ag can reduce it to bromide, but the potential for iodine is too high for Ag to reduce iodine to iodide.

What is the significance of silver's inability to reduce I_2 in practical applications?

It highlights silver's selective reactivity, making it useful in processes where bromine reduction is needed without affecting iodine compounds.

Is the reaction between Ag(s) and Br_2 spontaneous? How can you tell?

Yes, the reaction is spontaneous because silver readily forms silver bromide, indicating a favorable reduction process.

What other substances can serve as reducing agents for Br_2 but not for I_2?

Certain metals like copper or zinc can reduce bromine but may not be effective in reducing iodine under similar conditions.

Why is the difference in reactivity between Br_2 and I_2 important in chemical reactions involving silver?

Because it determines the selectivity of silver in forming specific halide compounds, influencing the outcome of synthesis and separation processes in chemistry.

Additional Resources

Understanding the Redox Behavior of Bromine and Iodine: A Detailed Analysis

Redox reactions are fundamental to inorganic chemistry, involving the transfer of electrons between species. The ability of a substance to act as a reducing or oxidizing agent depends on its position in the electrochemical series, standard reduction potentials, and the specific chemical environment. In this discussion, we focus on identifying a substance that can reduce bromine in its molecular form (Br_2) to bromide ions (Br^-) in aqueous solution but cannot reduce iodine (I_2) to iodide (I^-). This nuanced question probes the interplay of standard reduction potentials, chemical reactivity, and the nature of the oxidizing and reducing agents involved.

Fundamentals of Redox Chemistry: Standard Reduction Potentials and Their Significance

What Are Standard Reduction Potentials?

Standard reduction potentials (E°) are measures of the tendency of a species to gain electrons and be reduced under standard conditions (1 M concentration, 1 atm pressure, 25°C). They are tabulated in electrochemical series and serve as a basis for predicting the spontaneity of redox reactions.

- Positive E° indicates a high tendency to be reduced.
- Negative E° indicates a tendency to be oxidized.

Electrochemical Series for Halogens

Halogens exhibit a well-established order based on their reduction potentials:

Species	Reduction Reaction	E° (V)
F_2	$F_2 + 2e^- \rightarrow 2F^-$	+2.87
Cl_2	$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
Br_2	$Br_2 + 2e^- \rightarrow 2Br^-$	+1.07
I_2	$I_2 + 2e^- \rightarrow 2I^-$	+0.54

From this, we observe:
- Bromine (Br_2) is a stronger oxidant than iodine (I_2), given its higher reduction potential.
- The reduction potentials suggest that Br_2 can oxidize I^- to I_2, but the reverse reaction (reducing Br_2 to Br^-) is also feasible under suitable conditions.

Redox Behavior of Bromine and Iodine

Reduction of Bromine to Bromide

The reduction of bromine (Br_2) to bromide (Br^-) in aqueous solution is straightforward:

$$\text{Br}_2 (l) + 2e^- \rightarrow 2 \text{Br}^- \quad \text{(E° = +1.07 V)}$$

Any reducing agent with an oxidation potential lower than +1.07 V can, in principle, reduce Br_2 to Br^-.

Reduction of Iodine to Iodide

Similarly, iodine can be reduced as:

$$\text{I}_2 (s) + 2e^- \rightarrow 2 \text{I}^- \quad \text{(E° = +0.54 V)}$$

The reduction potential here is lower, indicating that I_2 is a weaker oxidant than Br_2 and easier to reduce.

Implication for Reducing Agents

- A substance capable of reducing Br_2 to Br^- must have an oxidation potential less than +1.07 V.
- To also reduce I_2 to I^-, the same agent must have an oxidation potential less than +0.54 V.

Therefore, the key point is that a reducing agent with an oxidation potential between +0.54 V and +1.07 V can reduce Br_2 to Br^- but cannot reduce I_2 to I^-, because it is not strong enough to reduce iodine.

Identifying a Suitable Substance: The Core Concept

Given the above, the core idea is to find a substance with an oxidation potential between +0.54 V and +1.07 V—strong enough to reduce bromine but not iodine.

Key points:
- The substance must be a good enough reducing agent to reduce Br_2 to Br^-.
- It must not be strong enough to reduce I_2 to I^-.

This criterion narrows down potential candidates significantly.

Potential Candidates and Their Redox Properties

Common Reducing Agents in the Electrochemical Series

Let's briefly examine some common reducing agents:

- Metallic Zinc (Zn): E° (Zn^{2+}/Zn) = −0.76 V
- Aluminum (Al): E° (Al^{3+}/Al) = −1.66 V
- Iron (Fe): E° (Fe^{3+}/Fe^{2+}) = +0.77 V, (Fe^{2+}/Fe) = −0.44 V
- Copper (Cu): E° (Cu^{2+}/Cu) = +0.34 V
- Hydrogen gas (H_2): E° ($2H^+ + 2e^- \rightleftharpoons H_2$) = 0 V

Among these, metals like zinc and iron are promising candidates because their oxidation potentials fall within the desired range.

Analysis:
- Zinc, with an oxidation potential of −0.76 V, can reduce both Br_2 and I_2.
- Copper (+0.34 V) can reduce I_2 but may not be sufficient to reduce Br_2.
- Iron (+0.77 V) can potentially reduce Br_2 but not I_2, since its potential is close but slightly higher than the reduction potential of I_2.

However, the actual redox behavior depends on kinetic factors, complex formation, and environmental conditions, so electrochemical potentials provide a theoretical basis rather than definitive proof.

Theoretical Candidate: The Use of Silver (Ag)

Standard Reduction Potential for Silver

- $Ag^+ + e^- \rightleftharpoons Ag(s)$: E° = +0.80 V

This potential suggests that silver metal can act as a reducing agent for species with reduction potentials less than +0.80 V, but not greater.

Implication:
- Silver can reduce I_2 (E° = +0.54 V) to I^-.
- It may not reduce Br_2 to Br^- because Br_2 has a higher reduction potential (+1.07 V), which is beyond Ag's

capability.

But, this is complicated because different factors, such as complex formation (like AgI or AgBr), can influence the actual reduction process.

Practical Identification: Using Silver as the Reducing Agent

Based on the electrochemical series, silver (Ag) is a promising candidate because:

- It can reduce I_2 to I^-:

$$ \text{Ag} (s) + \text{I}_2 (s) \rightarrow \text{AgI} (s) + I^- $$

because the reduction potential of I_2 is less than Ag's oxidation potential.

- It cannot reduce Br_2 to Br^-:

$$ \text{Ag} (s) + \text{Br}_2 (l) \rightarrow \text{AgBr} (s) + Br^- $$

is less favorable because Br_2's reduction potential (+1.07 V) exceeds Ag's oxidation potential (+0.80 V).

Therefore, silver can reduce iodine to iodide but cannot reduce bromine to bromide under standard conditions.

Additional Considerations and Experimental Evidence

Complex Formation and Kinetic Factors

While standard potentials give a thermodynamic perspective, actual reactions depend on:

- Kinetics: Reaction rates can be slow or fast depending on catalysts and environmental factors.
- Complexation: Silver forms insoluble halide precipitates (AgI, AgBr), which can drive the reactions forward or hinder them.

Silver halide formation:
- Silver iodide (AgI) is less soluble than silver bromide (AgBr).
- This difference influences the likelihood of reduction; AgI forms more readily due to lower solubility.

Implication in Practical Scenarios

- Silver can act as a selective reducing agent based on the halogen involved.
- Under controlled conditions, silver can reduce iodine but not bromine, making it an ideal candidate fitting the question's criteria.

Summary and Final Conclusion

The substance that can reduce Br_2 to Br^- but cannot reduce I_2 to I^- is best identified as:

Silver (Ag)

Rationale:
- Standard reduction potentials position silver as capable of reducing iodine ($E° = +0.54$ V) but not bromine ($E° = +1.07$ V).
- Chemical behavior supports that silver readily forms insoluble silver iodide (AgI), indicating reduction of iodine.
- Silver's oxidation potential (around +0.80 V) is insufficient to reduce bromine, preventing the formation of silver bromide under standard conditions.

Additional notes:
- The actual reduction behavior can vary with conditions such as pH, concentration, and presence of complexing agents.
- Silver's

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