

One Element, Two Opposite Changes

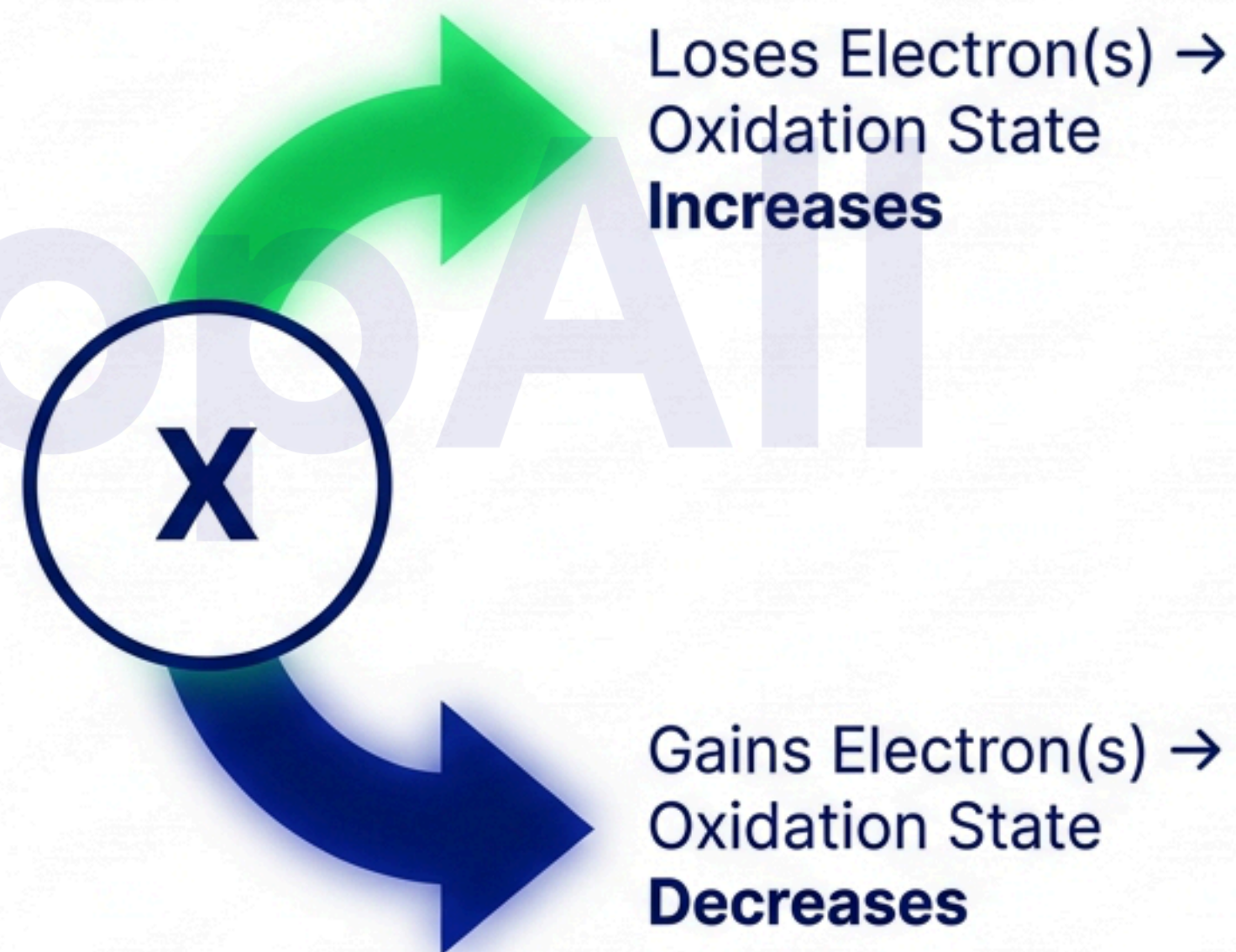
Understanding Disproportionation Reactions



A Special Type of Redox Reaction

Simultaneous oxidation and reduction of the exact same element.

A disproportionation reaction occurs when the **same element** in a reacting substance is **simultaneously oxidised and reduced**.



The Golden Requirement

Why an element must start in an intermediate oxidation state.



If an element is already in its highest or lowest possible oxidation state, it ***cannot*** disproportionate. It must have room to go both up and down.



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$$E=mc^2$$



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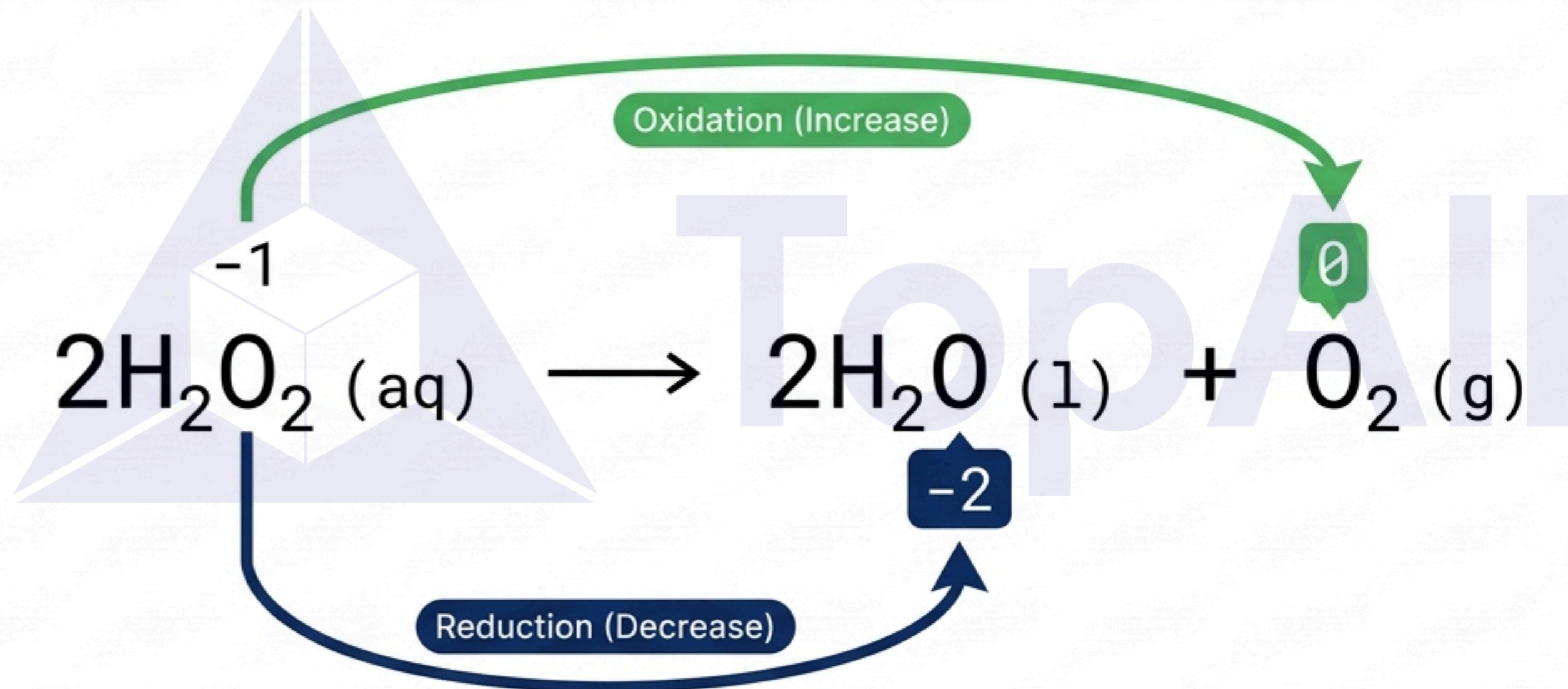
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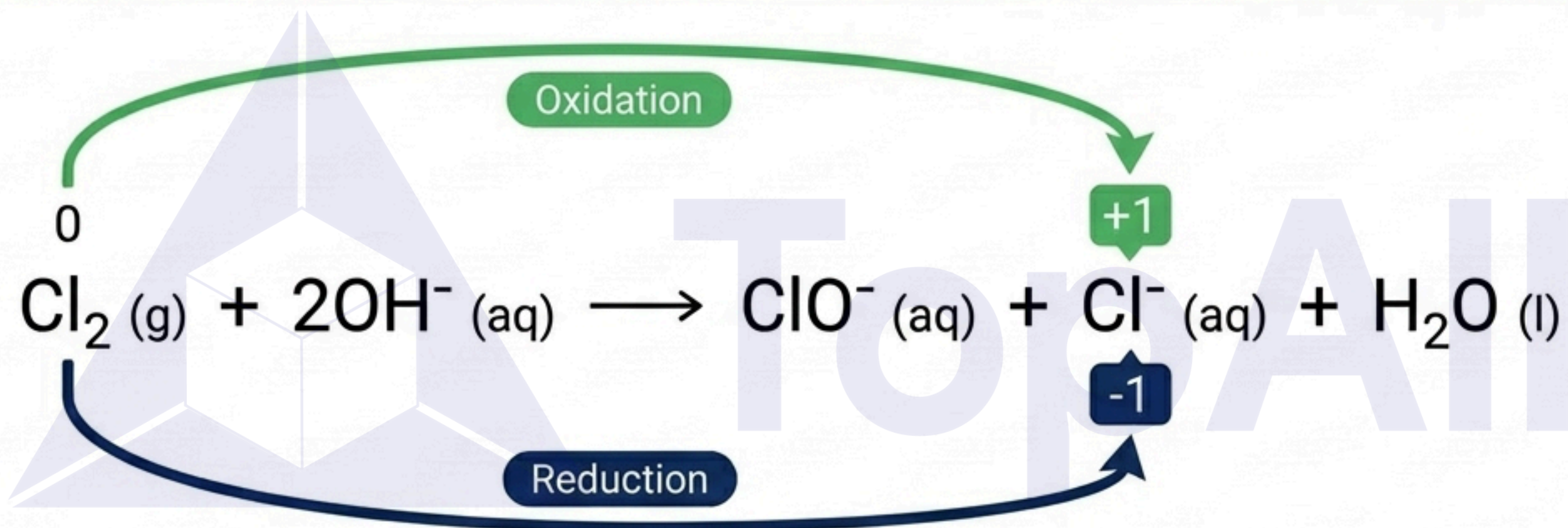
Classic Example: Hydrogen Peroxide

Oxygen acts as its own oxidant and reductant.



Industrial Example: **Chlorine in Alkali**

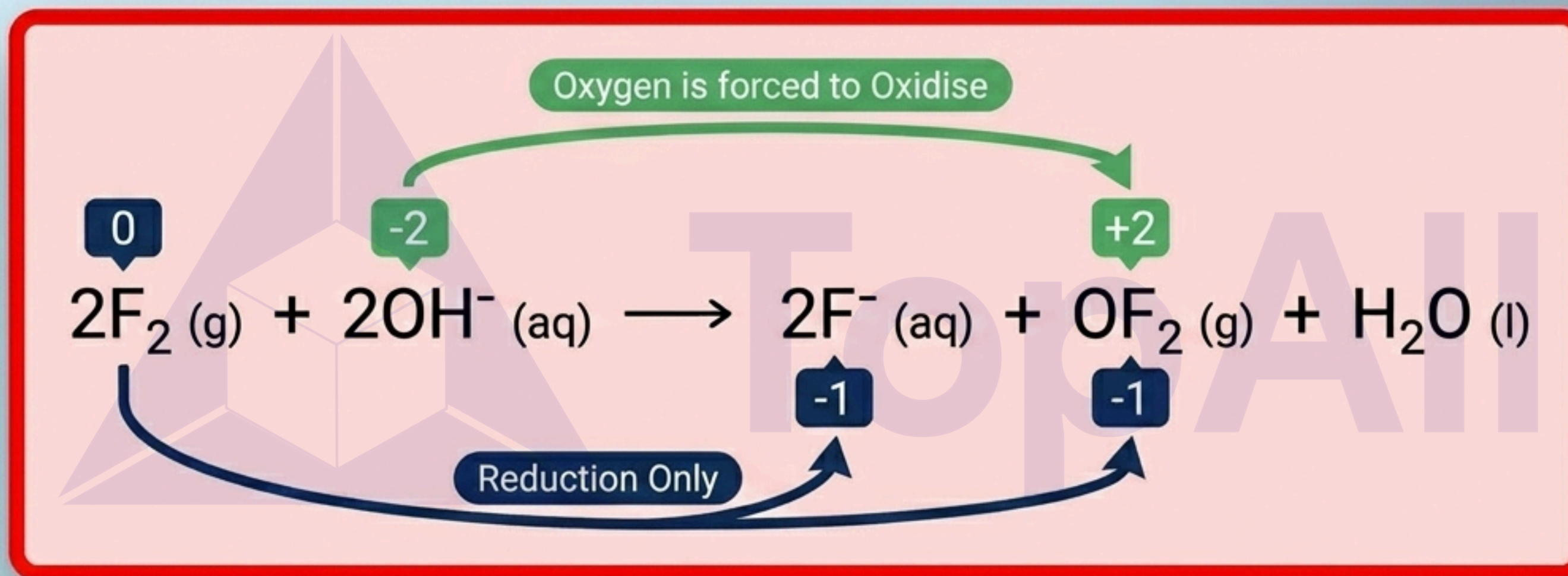
The chemistry behind household bleach.



The hypochlorite ion (ClO^-) formed here is the active ingredient in bleach. It oxidises color-bearing stains, turning them colorless.

The Fluorine Exception

Why the most electronegative element breaks the rule.



Fluorine is the most electronegative element. Its highest possible oxidation state is 0. Because it cannot exhibit a positive oxidation state, it **cannot be oxidised**. Therefore, Fluorine **NEVER** disproportionates.

Disproportionation Summary

NEET Mastery Checklist

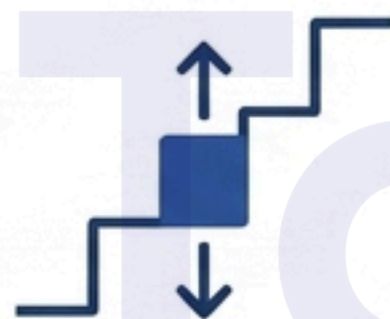


The Definition



One element splits into two paths simultaneously: oxidized (state increases) and reduced (state decreases).

The Requirement



The reactant element **MUST** be in an intermediate oxidation state. It needs room to move up and down.

The Trap



Fluorine **NEVER** disproportionates. As the most electronegative element, it cannot achieve a positive oxidation state.

Review Complete



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