RedOx Chemistry

When it's barely chemistry, it's RedOx Chemistry

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What is Chemistry?

Chemistry is often defined as "making and breaking bonds"; rearranging atoms to form new substances.

There is one class of molecular reactions that is incredibly important but defies this definition: electrochemistry.

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Consider 2 molecules

FeO and Fe₂O₃

Are they different?
Yes.

What's the difference?
Iron (II) oxide vs. Iron (III) oxide  The Oxidation State is different.
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Are you stuck with your oxidation state?

Asked a different way: If you are iron in FeO, are you stuck being Fe$^{2+}$ forever?

In fact, you can change oxidation states as often as you like. But, there’s a catch...

How do you change oxidation states?

Add or subtract electrons. Fe$^{2+}$ has 1 more electron than Fe$^{3+}$.

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What does this reaction look like?

Fe$^{2+} \rightarrow Fe^{3+} + 1 e^-$

Is this a “real” reaction?

Depends on what you mean by “real” and by reaction. Something changed, but no atoms were rearranged so it isn’t like the other reactions we’ve seen before. And, you might ask, what happens to the electron?

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This is an “electrochemical” reaction

Fe$^{2+} \rightarrow Fe^{3+} + 1 e^-$

It’s a special kind of process, part electrical and part (barely) chemical. The atom changes oxidation state and creates an electron. The electron can do useful work (power your iPod) or chemical work (change the oxidation state of something else).
Electrons come, electrons go

Fe\(^{2+}\) \rightarrow Fe^{3+} + 1 e^{-}
Mn^{5+} + 3 e^{-} \rightarrow Mn^{2+}

When electrons "go", it is called an "oxidation". When electrons "come", it is called a "reduction".

[It's easiest to remember that a "reduction" reduces the charge on the ion (oxidation state).]

Like acids and bases…

Oxidation and Reduction always happens simultaneously:

Oxidation half-reaction: Fe\(^{2+}\) \rightarrow Fe^{3+} + 1 e^{-}
Reduction half-reaction: Mn^{5+} + 3 e^{-} \rightarrow Mn^{2+}

Full reaction: 3 Fe\(^{2+}\) + Mn\(^{5+}\) \rightarrow 3 Fe^{3+} + Mn^{2+}

WTFDYGT?????????????

Chemical reactions don’t have electrons

Oxidation and Reduction half-reactions balance so that no NET electrons remain

Oxidation gives you 1 e-: Fe\(^{2+}\) \rightarrow Fe^{3+} + 1 e^{-}
Reduction needs 3: Mn\(^{5+}\) + 3 e^{-} \rightarrow Mn^{2+}

3 x (Fe\(^{2+}\) \rightarrow Fe^{3+} + 1 e^{-})
+ Mn\(^{5+}\) + 3 e^{-} \rightarrow Mn^{2+}
3 Fe\(^{2+}\) + Mn\(^{5+}\) + 3 e^{-} \rightarrow 3 Fe^{3+} + Mn^{2+}
3 Fe\(^{2+}\) + Mn\(^{5+}\) \rightarrow 3 Fe^{3+} + Mn^{2+}
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Is it always that easy?

Of course NOT!

Unbalanced equation:

\[ \text{CuO} + \text{FeO} \rightarrow \text{Fe}_2\text{O}_3 + \text{Cu}_2\text{O} \]

What’s going on here?

Well, it is a redox reaction but it is a little less obvious than when I am just showing the ions. The oxidation state is hidden in the molecules.

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Is it always that easy?

\[ \text{CuO} + \text{FeO} \rightarrow \text{Fe}_2\text{O}_3 + \text{Cu}_2\text{O} \]

\[ \text{CuO} \rightarrow \text{copper (II) oxide} \]

\[ \text{Cu}_2\text{O} \rightarrow \text{copper (I) oxide} \]

\[ \text{FeO} \rightarrow \text{iron (II) oxide} \]

\[ \text{Fe}_2\text{O}_3 \rightarrow \text{iron (III) oxide} \]

How do you know? Remember our nomenclature: O is always -2, halogens are -1, etc.

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Is it always that easy?

\[ \text{CuO} + \text{FeO} \rightarrow \text{Fe}_2\text{O}_3 + \text{Cu}_2\text{O} \]

\[ \text{CuO} \rightarrow \text{copper (II) oxide} \]

\[ \text{Cu}_2\text{O} \rightarrow \text{copper (I) oxide} \]

\[ \text{FeO} \rightarrow \text{iron (II) oxide} \]

\[ \text{Fe}_2\text{O}_3 \rightarrow \text{iron (III) oxide} \]

Looked at this way, it is clearer that the Cu is going from +2 on the left to +1 on the right (reduction) at the same time that the iron is going from +2 on the left to +3 on the right (oxidation).
How do I balance the equation?

CuO + FeO \rightarrow Fe_2O_3 + Cu_2O

Balancing redox reactions is similar to regular equations BUT it also requires that you balance the charges as well.

Fortunately, there is a relatively easy system that ALWAYS works! Just follow the 7-ish easy steps!

1. Separate into ½ reactions

CuO + FeO \rightarrow Fe_2O_3 + Cu_2O

Break the full reaction into 2 half-reactions:

Oxidation: FeO \rightarrow Fe_2O_3

Reduction: CuO \rightarrow Cu_2O

We treat them separately from now on.

2. Balance each ½ reaction, ignoring O and H

Oxidation: 2 FeO \rightarrow Fe_2O_3

Reduction: 2 CuO \rightarrow Cu_2O

Just want same number of atoms on each side.
3 – Balance the oxygen by adding water

This is more logical than it seems since most electrochemistry occurs in aqueous media.

Oxidation: $2 \text{FeO} \rightarrow \text{Fe}_2\text{O}_3$

Reduction: $2 \text{CuO} \rightarrow \text{Cu}_2\text{O}$

Oxidation: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3$

Reduction: $2 \text{CuO} \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$

4 – Balance the hydrogen by adding $\text{H}^+$

This is also more logical than it seems, since aqueous solutions (as we’ve seen) are generally either acidic or basic.

Oxidation: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3$

Reduction: $2 \text{CuO} \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$

Oxidation: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2 \text{H}^+$

Reduction: $2 \text{CuO} + 2 \text{H}^+ \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$

The atoms are balanced

At this point, the two half-reactions should be balanced based only on the atoms. But notice that the charge isn't balanced!

Oxidation: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2 \text{H}^+$

Reduction: $2 \text{CuO} + 2 \text{H}^+ \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$
5 – Balance the charges by adding electrons

Oxidation: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2 \text{H}^+$
- $0$ charge $\rightarrow +2$ charge

Reduction: $2 \text{CuO} + 2 \text{H}^+ \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$
- $+2$ charge $\rightarrow 0$ charge

Oxidation: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2 \text{H}^+ + 2$ e$^-$$

Reduction: $2 \text{CuO} + 2 \text{H}^+ + 2$ e$^- \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$

ALWAYS add the electrons to the more POSITIVE side.

I’M NOT TRYING TO MAKE THE CHARGE ZERO!

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6 – Combine the half-reaction, eliminating any electrons

I want to add the 2 reactions together, making sure the electrons cancel on each side. (easy here)

Ox: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2 \text{H}^+ + 2$ e$^-$
- 2 electrons on right

Red: $2 \text{CuO} + 2 \text{H}^+ + 2$ e$^- \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$
- 2 electrons on left

I just add them together as is. If there were a different number of electrons, I’d need to multiply the reactions by whatever factors make them the same.

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6 – Combine the half-reaction, eliminating any electrons and canceling common components

I want to add the 2 reactions together, making sure the electrons cancel on each side. (easy here)

Ox: $2 \text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2 \text{H}^+ + 2$ e$^-$$

Red: $2 \text{CuO} + 2 \text{H}^+ + 2$ e$^- \rightarrow \text{Cu}_2\text{O} + \text{H}_2\text{O}$
- $2 \text{FeO} + 2 \text{CuO} + 2 \text{H}^+ + 2$ e$^- \rightarrow \text{Fe}_2\text{O}_3 + \text{Cu}_2\text{O} + 3$ H$^+$
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7-ish – IF in basic solution rather than acid, add OH⁻ to both sides to eliminate the H⁺

2 FeO + 2 CuO → Fe₂O₃ + Cu₂O

Not a factor here!

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New example:
Balance the following equation in basic solution:

ClO₄⁻(aq) + Cl⁻(aq) → ClO₃⁻(aq) + Cl₂(aq)

We just need to apply our 7-ish steps.

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ClO₄⁻(aq) + Cl⁻(aq) → ClO₃⁻(aq) + Cl₂(aq)

That reaction SCREAMS base because...
A. You seem to think it should.
B. Conjugate bases
C. Negative ions
D. None of the above (including E)
E. All of the above (well except D)
F. Happy Thanksgiving
New example:
Balance the following equation in basic solution:

\[ \text{ClO}_4^- (aq) + \text{Cl}^- (aq) \rightarrow \text{ClO}_3^- (aq) + \text{Cl}_2 (aq) \]

1 – Separate into \( \frac{1}{2} \) reactions

\[ \text{ClO}_4^- (aq) + \text{Cl}^- (aq) \rightarrow \text{ClO}_3^- (aq) + \text{Cl}_2 (aq) \]

What’s changing oxidation state?
\( \text{Cl}^- \) - oxidation state is -1
\( \text{Cl}_2 \) - oxidation state is 0 (all elementals are 0)
\( \text{ClO}_4^- \) - WTFITOS?
\( \text{ClO}_3^- \) - WTFITOS?

What’s changing oxidation state?
\( \text{Cl}^- \) - oxidation state is -1
\( \text{Cl}_2 \) - oxidation state is 0 (all elementals are 0)
\( \text{ClO}_4^- \) - Cl is +7 (O is -2, ion is -1 overall)
\( \text{ClO}_3^- \) - Cl is +5 (O is -2, ion is -1 overall)
1 – Separate into ½ reactions

\[ \text{ClO}_4^- (aq) + \text{Cl}^- (aq) \rightarrow \text{ClO}_3^- (aq) + \text{Cl}_2 (aq) \]

Break the full reaction into 2 half-reactions:

- Oxidation: \( \text{Cl}^- (aq) \rightarrow \text{Cl}_2 (aq) \)
- Reduction: \( \text{ClO}_4^- (aq) \rightarrow \text{ClO}_3^- (aq) \)

We treat them separately from now on.

2 – Balance each ½ reaction, ignoring O and H

- Oxidation: \( 2 \text{Cl}^- (aq) \rightarrow \text{Cl}_2 (aq) \)
- Reduction: \( \text{ClO}_4^- (aq) \rightarrow \text{ClO}_3^- (aq) \)

Just want same number of atoms on each side.

- Oxidation: \( 2 \text{Cl}^- (aq) \rightarrow \text{Cl}_2 (aq) \)
- Reduction: \( \text{ClO}_4^- (aq) \rightarrow \text{ClO}_3^- (aq) + \text{H}_2\text{O}(l) \)

3 – Balance the oxygen by adding water

- Oxidation: \( 2 \text{Cl}^- (aq) \rightarrow \text{Cl}_2 (aq) \)
- Reduction: \( \text{ClO}_4^- (aq) \rightarrow \text{ClO}_3^- (aq) \)

Oxidation: \( 2 \text{Cl}^- (aq) \rightarrow \text{Cl}_2 (aq) \)

Reduction: \( \text{ClO}_4^- (aq) \rightarrow \text{ClO}_3^- (aq) + \text{H}_2\text{O}(l) \)
4 – Balance the hydrogen by adding H+

Oxidation: 2 Cl\(^-\) (aq) \rightarrow Cl\(_2\) (aq)

Reduction: ClO\(_4\)\(^-\) (aq) \rightarrow ClO\(_3\)\(^-\) (aq) + H\(_2\)O(l)

Oxidation: 2 Cl\(^-\) (aq) \rightarrow Cl\(_2\) (aq)

Reduction: ClO\(_4\)\(^-\) + 2 H\(^+\) (aq) \rightarrow ClO\(_3\)\(^-\) (aq) + H\(_2\)O(l)

5 – Balance the charges by adding electrons

Oxidation: 2 Cl\(^-\) (aq) \rightarrow Cl\(_2\) (aq)

Reduction: ClO\(_4\)\(^-\) (aq) + 2 H\(^+\) (aq) \rightarrow ClO\(_3\)\(^-\) (aq) + H\(_2\)O(l)

2 Cl\(^-\) (aq) \rightarrow Cl\(_2\) (aq) + 2 e\(^-\)

Oxidation: 2 Cl\(^-\) (aq) \rightarrow Cl\(_2\) (aq) + 2 e\(^-\)

Reduction: ClO\(_4\)\(^-\) (aq) + 2 H\(^+\) (aq) + 2 e\(^-\) \rightarrow ClO\(_3\)\(^-\) (aq) + H\(_2\)O(l)

6 – Combine the half-reaction, eliminating any electrons

Ox: 2 Cl (aq) \rightarrow Cl\(_2\) (aq) + 2 e\(^-\)

Red: ClO\(_4\)\(^-\) (aq) + 2 H\(^+\) (aq) + 2 e\(^-\) \rightarrow ClO\(_3\)\(^-\) (aq) + H\(_2\)O(l)

2 Cl\(^-\) (aq) + ClO\(_4\)\(^-\) (aq) + 2 H\(^+\) (aq) + 2 e\(^-\) \rightarrow Cl\(_2\) (aq) + ClO\(_3\)\(^-\) (aq) + H\(_2\)O(l)

2 Cl\(^-\) (aq) + ClO\(_4\)\(^-\) (aq) + 2 H\(^+\) (aq) \rightarrow Cl\(_2\) (aq) + ClO\(_3\)\(^-\) (aq) + H\(_2\)O(l)
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7-ish – IF in basic solution rather than acid, add OH\(^-\) to both sides to eliminate the H\(^+\)

\[
\begin{align*}
2 \text{Cl}^- (aq) + \text{ClO}_4^- (aq) + 2 \text{H}^+ (aq) & \rightarrow \text{Cl}_2 (aq) + \text{ClO}_2^- (aq) + H_2O(l) \\
2 \text{Cl}^- (aq) + \text{ClO}_4^- (aq) + 2 \text{H}^+ (aq) & \rightarrow \text{Cl}_2 (aq) + \text{ClO}_3^- (aq) + 2 \text{H}^+(aq) + 2 \text{OH}^- (aq)
\end{align*}
\]

Why 2 OH\(^-\)? Because I need to neutralize 2 H\(^+\) which gives me... 2 H\(_2\)O!!

\[
\begin{align*}
2 \text{Cl}^- (aq) + \text{ClO}_4^- (aq) + 2 \text{H}^+(aq) + 2 \text{OH}^- (aq) & \rightarrow \text{Cl}_2 (aq) + \text{ClO}_3^- (aq) + \text{H}_2\text{O}(l) + 2 \text{OH}^- (aq)
\end{align*}
\]

Cleaning up a little bit:

\[
\begin{align*}
2 \text{Cl}^- (aq) + \text{ClO}_4^- (aq) + 2 \text{H}_2\text{O}(l) & \rightarrow \text{Cl}_2 (aq) + \text{ClO}_3^- (aq) + \text{H}_2\text{O}(l) + 2 \text{OH}^- (aq)
\end{align*}
\]

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7-ish – IF in basic solution rather than acid, add OH\(^-\) to both sides to eliminate the H\(^+\)

Cleaning up a little bit:

\[
\begin{align*}
2 \text{Cl}^- (aq) + \text{ClO}_4^- (aq) + 2 \text{H}_2\text{O}(l) & \rightarrow \text{Cl}_2 (aq) + \text{ClO}_3^- (aq) + \text{H}_2\text{O}(l) + 2 \text{OH}^- (aq)
\end{align*}
\]

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One more example:

Balance the following equation in basic solution:

\[
\text{I}^- (aq) + \text{NO}_2^- (aq) \rightarrow \text{I}_2 (s) + \text{NO} (g)
\]
1 – Separate into ½ reactions

\[ I^- (aq) + NO_2^- (aq) \rightarrow I_2 (s) + NO (g) \]

What's changing oxidation state?

- I\(^-\): oxidation state is -1
- \(I_2\): oxidation state is 0 (all elements are 0)
- \(NO_2^-\): N is +3
- NO: N is +2

Break the full reaction into 2 half-reactions:

Oxidation: \(I^- (aq) \rightarrow I_2 (s)\)

Reduction: \(NO_2^- (aq) \rightarrow NO (g)\)
2 – Balance each ½ reaction, ignoring O and H

Oxidation: I\(^{-}\)\(_{(aq)}\) → I\(_2\)\(_{(s)}\)
Reduction: NO\(_2\)\(^{-}\)\(_{(aq)}\) → NO\(_{(g)}\)

Just want the same number of atoms on each side.

Oxidation: \(2 \ I\(^{-}\)\(_{(aq)}\) → 1 \ I\(_2\)\(_{(s)}\)\)
Reduction: \(\text{NO}_2\)\(^{-}\)\(_{(aq)}\) → NO\(_{(g)}\) + H\(_2\)O\(_{(l)}\)

3 – Balance the oxygen by adding water

Oxidation: \(2 \ I\ (^{-})_{(aq)} \rightarrow I_2 \,(s)\)
Reduction: \(\text{NO}_2\)\(^{-}\)\(_{(aq)}\) → NO\(_{(g)}\) + H\(_2\)O\(_{(l)}\)

Oxidation: \(2 \ I\ (^{-})_{(aq)} \rightarrow I_2 \,(s)\)
Reduction: \(\text{NO}_2\)\(^{-}\)\(_{(aq)}\) → NO\(_{(g)}\) + H\(_2\)O\(_{(l)}\)

4 – Balance the hydrogen by adding H\(^+\)

Oxidation: \(2 \ I\ (^{-})_{(aq)} \rightarrow I_2 \,(s)\)
Reduction: \(\text{NO}_2\)\(^{-}\)\(_{(aq)}\) → NO\(_{(g)}\) + H\(_2\)O\(_{(l)}\)

Oxidation: \(2 \ I\ (^{-})_{(aq)} \rightarrow I_2 \,(s)\)
Reduction: \(\text{NO}_2\)\(^{-}\)\(_{(aq)}\) + 2 \ H\(^+\)\(_{(aq)}\) → NO\(_{(g)}\) + H\(_2\)O\(_{(l)}\)
5 – Balance the charges by adding electrons

Ox: \( 2 \text{I}^- (\text{aq}) \rightarrow \text{I}_2 (\text{s}) \)

Red: \( \text{NO}_2^- (\text{aq}) + 2 \text{H}^+ (\text{aq}) \rightarrow \text{NO} (\text{g}) + \text{H}_2\text{O} (\text{l}) \)

Ox: \( 2 \text{I}^- (\text{aq}) \rightarrow \text{I}_2 (\text{s}) + 2 \text{e}^- \)

Red: \( \text{NO}_2^- (\text{aq}) + 2 \text{H}^+ (\text{aq}) + 1 \text{e}^- \rightarrow \text{NO} (\text{g}) + \text{H}_2\text{O} (\text{l}) \)

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6 – Combine the half-reaction, eliminating any electrons

Ox: \( 2 \text{I}^- (\text{aq}) \rightarrow \text{I}_2 (\text{s}) + 2 \text{e}^- \)

Red: \( \text{NO}_2^- (\text{aq}) + 2 \text{H}^+ (\text{aq}) + 1 \text{e}^- \rightarrow \text{NO} (\text{g}) + \text{H}_2\text{O} (\text{l}) \)

Ox + 2*Red

Ox: \( 2 \text{I}^- (\text{aq}) \rightarrow \text{I}_2 (\text{s}) + 2 \text{e}^- \)

Red: \( 2 \text{NO}_2^- (\text{aq}) + 4 \text{H}^+ (\text{aq}) + 2 \text{e}^- \rightarrow 2 \text{NO} (\text{g}) + 2 \text{H}_2\text{O} (\text{l}) \)
6 – Combine the half-reaction, eliminating any electrons

Ox: \(2 \text{I}^- (\text{aq}) \rightarrow \text{I}_2 (s) + 2 e^-\)

Red: \(2 \text{NO}_2^- (\text{aq}) + 4 \text{H}^+ (\text{aq}) + 2 e^- \rightarrow 2 \text{NO} (g) + 2 \text{H}_2\text{O} (l)\)

\[
2 \text{I}^- (\text{aq}) + 2 \text{NO}_2^- (\text{aq}) + 4 \text{H}^+ (\text{aq}) + 2 e^- \rightarrow \text{I}_2 (s) + 2 \text{NO} (g) + 2 \text{H}_2\text{O} (l)
\]

7-ish – IF in basic solution rather than acid, add OH\(^-\) to both sides to eliminate the H\(^+\)

Cleaning up a little bit:

\[
2 \text{I}^- (\text{aq}) + 2 \text{NO}_2^- (\text{aq}) + 4 \text{H}^+ (\text{aq}) + 4 \text{OH}^- \rightarrow \text{I}_2 (s) + 2 \text{NO} (g) + 4 \text{H}_2\text{O} (l) + 4 \text{OH}^-\]

Why 4 OH\(^-\)? Because I need to neutralize 4 H\(^+\) which gives me...4 H\(_2\)O!!

\[
2 \text{I}^- (\text{aq}) + 2 \text{NO}_2^- (\text{aq}) + 4 \text{H}_2\text{O} (l) \rightarrow \text{I}_2 (s) + 2 \text{NO} (g) + 4 \text{OH}^-\]
\]
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6-1/2 Magic steps

1. Separate into ½ reactions
2. Balance ½ reactions except for O, H
3. Balance O by adding H2O
4. Balance H by adding H+ 
5. Balance charge by adding electrons
6. Combine ½ reactions, eliminating electrons as you do it.
7. IF IF IF in basic solution, "neutralize" the H+ by adding OH- to both sides

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I make an electrochemical cell by mixing 2 half reactions in acidic solutions:

O2(g) → H2O (l)

H2SO3 (aq) → SO4

2-
(aq)

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O2(g) → H2O (l) + H2O

H2SO3 (aq) + H2O → SO4

2-
(aq)
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\[ \text{O}_2(\text{g}) + 4\text{H}^+ \rightarrow \text{H}_2\text{O} (\text{l}) + \text{H}_2\text{O} \]

\[ \text{H}_2\text{SO}_3 (\text{aq}) + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} (\text{aq}) + 4 \text{H}^+ \]

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\[ \text{O}_2(\text{g}) + 4\text{H}^+ + 4 \text{e}^- \rightarrow \text{H}_2\text{O} (\text{l}) + \text{H}_2\text{O} \]

\[ \text{H}_2\text{SO}_3 (\text{aq}) + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 4 \text{H}^+ + 2\text{e}^- \]

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\[ \text{O}_2(\text{g}) + 4\text{H}^+ + 4 \text{e}^- \rightarrow \text{H}_2\text{O} (\text{l}) + \text{H}_2\text{O} \]

\[ 2\times[\text{H}_2\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^-] \]

\[ 2\text{H}_2\text{SO}_3 + 2 \text{H}_2\text{O} \rightarrow 2\text{SO}_4^{2-} + 8\text{H}^+ + 4\text{e}^- \]
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\[ \text{O}_2(g) + 4\text{H}^+ + 4 \text{e}^- \rightarrow \text{H}_2\text{O} (l) + \text{H}_2\text{O} \]

\[ 2\text{H}_2\text{SO}_3 + 2 \text{H}_2\text{O} \rightarrow 2\text{SO}_4^{2-} + 8\text{H}^+ + 4\text{e}^- \]

\[ 2\text{H}_2\text{SO}_3 + 2 \text{H}_2\text{O} \rightarrow 2\text{SO}_4^{2-} + 8\text{H}^+ + 4\text{e}^- \]

\[ \text{O}_2(g) + 2\text{H}_2\text{SO}_3 \rightarrow 2\text{SO}_4^{2-} + 4\text{H}^+ \]

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\[ \text{O}_2(g) + 4\text{H}^+ + 4 \text{e}^- \rightarrow \text{H}_2\text{O} (l) + \text{H}_2\text{O} \]

\[ 2\text{H}_2\text{SO}_3 + 2 \text{H}_2\text{O} \rightarrow 2\text{SO}_4^{2-} + 8\text{H}^+ + 4\text{e}^- \]

\[ 2\text{H}_2\text{SO}_3 + 2 \text{H}_2\text{O} \rightarrow 2\text{SO}_4^{2-} + 8\text{H}^+ + 4\text{e}^- \]

\[ \text{O}_2(g) + 2\text{H}_2\text{SO}_3 \rightarrow 2\text{SO}_4^{2-} + 4\text{H}^+ \]