Slide 1

RedOx Chemistry

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

Slide 2

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 Fe_2O_3

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 is 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).

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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).]

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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?????????????

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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: $Mn^{5+} + 3\ e- \rightarrow Mn^{2+}$

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

$+ \underline{Mn^{5+} + 3\ e- \rightarrow Mn^{2+}}$

$3\ Fe^{2+} + Mn^{5+} + \cancel{3e^-} \rightarrow 3\ Fe^{3+} + Mn^{2+} + \cancel{3e^-}$

$3\ Fe^{2+} + Mn^{5+} \rightarrow 3\ Fe^{3+} + Mn^{2+}$

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Is it always that easy?

Of course NOT!
Unbalanced equation:
$CuO + FeO \rightarrow Fe_2O_3 + Cu_2O$

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?

$CuO + FeO \rightarrow Fe_2O_3 + Cu_2O$

CuO – copper (II) oxide
Cu_2O – copper (I) oxide

FeO – iron (II) oxide
Fe_2O_3 – 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?

$CuO + FeO \rightarrow Fe_2O_3 + Cu_2O$

CuO – copper (II) oxide
Cu_2O – copper (I) oxide

FeO – iron (II) oxide
Fe_2O_3 – 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).

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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!

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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.

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2 – Balance each ½ reaction, ignoring O and H

Oxidation: $FeO \rightarrow Fe_2O_3$

Reduction: $CuO \rightarrow Cu_2O$

Just want same number of atoms on each side.

Oxidation: $2\ FeO \rightarrow Fe_2O_3$

Reduction: $2\ CuO \rightarrow Cu_2O$

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3 – Balance the oxygen by adding water

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

Oxidation: $2\ FeO \rightarrow Fe_2O_3$
2 O 3 O
Reduction: $2\ CuO \rightarrow Cu_2O$
2 O 1 O
Oxidation: $2\ FeO + H_2O \rightarrow Fe_2O_3$

Reduction: $2\ CuO \rightarrow Cu_2O + H_2O$

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4 – Balance the hydrogen by adding H^+

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

Oxidation: $2\ FeO + H_2O \rightarrow Fe_2O_3$

Reduction: $2\ CuO \rightarrow Cu_2O + H_2O$

Oxidation: $2\ FeO + H_2O \rightarrow Fe_2O_3 + 2\ H^+$

Reduction: $2\ CuO + 2\ H^+ \rightarrow Cu_2O + H_2O$

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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: $2\ FeO + H_2O \rightarrow Fe_2O_3 + 2\ H^+$
0 charge +2 charge
Reduction: $2\ CuO + 2\ H^+ \rightarrow Cu_2O + H_2O$
+2 charge 0 charge

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5 – Balance the charges by adding electrons

Oxidation: $2\ FeO + H_2O \rightarrow Fe_2O_3 + 2\ H^+$

0 charge +2 charge

Reduction: $2\ CuO + 2\ H^+ \rightarrow Cu_2O + H_2O$

+2 charge 0 charge

Oxidation: $2\ FeO + H_2O \rightarrow Fe_2O_3 + 2\ H^+ + 2\ e^-$

Reduction: $2\ CuO + 2\ H^+ + 2e^- \rightarrow Cu_2O + H_2O$

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: $2\ FeO + H_2O \rightarrow Fe_2O_3 + 2\ H^+ + 2\ e^-$

2 electrons on right

Red: $2\ CuO + 2\ H^+ + 2e^- \rightarrow Cu_2O + H_2O$

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: $2\ FeO + H_2O \rightarrow Fe_2O_3 + 2\ H^+ + 2\ e^-$

Red: $\underline{2\ CuO + 2\ H^+ + 2e^- \rightarrow Cu_2O + H_2O}$

$2\ FeO + \cancel{H_2O} + 2\ CuO + \cancel{2\ H^+} + \cancel{2e^-} \rightarrow$
$Fe_2O_3 + \cancel{2\ H^+} + \cancel{2\ e^-} + Cu_2O + \cancel{H_2O}$

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

$2\ FeO + 2\ CuO \rightarrow Fe_2O_3 + Cu_2O$

Not a factor here!

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New example:

Balance the following equation in basic solution:

$ClO_4^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2\,(aq)}$

We just need to apply our 7-ish steps.

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$ClO_4^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2\,(aq)}$

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

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New example:

Balance the following equation in basic solution:

$ClO_4^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2\,(aq)}$

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1 – Separate into ½ reactions

$ClO_4^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2\,(aq)}$

What's changing oxidation state?

Cl^- - oxidation state is -1

Cl_2 - oxidation state is 0 (all elementals are 0)

ClO_4^- WTFITOS?

ClO_3^- WTFITOS?

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1 – Separate into ½ reactions

$ClO_4^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2\,(aq)}$

What's changing oxidation state?

Cl^- - oxidation state is -1

Cl_2 - oxidation state is 0 (all elementals are 0)

ClO_4^- - Cl is +7 (O is -2, ion is -1 overall)

ClO_3^- - Cl is +5 (O is -2, ion is -1 overall)

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1 – Separate into ½ reactions

$ClO_4^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + Cl_{2\,(aq)}$

Break the full reaction into 2 half-reactions:

Oxidation: $Cl^-{}_{(aq)} \rightarrow Cl_{2\,(aq)}$

Reduction: $ClO_4^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)}$

We treat them separately from now on.

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2 – Balance each ½ reaction, ignoring O and H

Oxidation: $Cl^-{}_{(aq)} \rightarrow Cl_{2\,(aq)}$

Reduction: $ClO_4^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)}$

Just want same number of atoms on each side.

Oxidation: $2\ Cl^-{}_{(aq)} \rightarrow Cl_{2\,(aq)}$

Reduction: $ClO_4^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)}$

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3 – Balance the oxygen by adding water

Oxidation: $2\ Cl^-{}_{(aq)} \rightarrow Cl_{2\,(aq)}$

Reduction: $ClO_4^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)}$

Oxidation: $2\ Cl^-{}_{(aq)} \rightarrow Cl_{2\,(aq)}$

Reduction: $ClO_4^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + H_2O_{(l)}$

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4 – Balance the hydrogen by adding H^+

Oxidation: $2\ Cl^-_{(aq)} \rightarrow Cl_{2\ (aq)}$

Reduction: $ClO_4^-{}_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + H_2O_{(l)}$

Oxidation: $2\ Cl^-_{(aq)} \rightarrow Cl_{2\ (aq)}$

Reduction: $ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + H_2O_{(l)}$

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5 – Balance the charges by adding electrons

Oxidation: $2\ Cl^-_{(aq)} \rightarrow Cl_{2\ (aq)}$
2*(-1)   0

Reduction: $ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + H_2O_{(l)}$
-1 + 2(1+) =+1   -1

Oxidation: $2\ Cl^-_{(aq)} \rightarrow Cl_{2\ (aq)} + 2\ e^-$

Reduction: $ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} + 2\ e^- \rightarrow ClO_3^-{}_{(aq)} + H_2O_{(l)}$

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6 – Combine the half-reaction, eliminating any electrons

Ox: $2\ Cl^-_{(aq)} \rightarrow Cl_{2\ (aq)} + 2\ e^-$
2 electrons

Red: $ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} + 2\ e^- \rightarrow ClO_3^-{}_{(aq)} + H_2O_{(l)}$
2 electrons

$2\ Cl^-_{(aq)} + ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} + \cancel{2\ e^-} \rightarrow Cl_{2\ (aq)} + \cancel{2\ e^-} + ClO_3^-{}_{(aq)} + H_2O_{(l)}$

$2\ Cl^-_{(aq)} + ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} \rightarrow Cl_{2\ (aq)} + ClO_3^-{}_{(aq)} + H_2O_{(l)}$

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

$2\ Cl^-_{(aq)} + ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} \rightarrow Cl_{2\ (aq)} + ClO_3^-{}_{(aq)} + H_2O_{(l)}$

$2\ Cl^-_{(aq)} + ClO_4^-{}_{(aq)} + 2\ H^+_{(aq)} + 2\ OH^- \rightarrow Cl_{2\ (aq)} + ClO_3^-{}_{(aq)} + H_2O_{(l)} + 2\ OH^-$

Why 2 OH-? Because I need to neutralize 2 H^+ which gives me…2 H_2O!!

$2\ Cl^-_{(aq)} + ClO_4^-{}_{(aq)} + 2\ H_2O_{(l)} \rightarrow Cl_{2\ (aq)} + ClO_3^-{}_{(aq)} + H_2O_{(l)} + 2\ OH^-$

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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:

$2\ Cl^-_{(aq)} + ClO_4^-{}_{(aq)} + \cancel{2}\ H_2O_{(l)} \rightarrow Cl_{2\ (aq)} + ClO_3^-{}_{(aq)} + \cancel{H_2O_{(l)}} + 2\ OH^-$

$2\ Cl^-_{(aq)} + ClO_4^-{}_{(aq)} + H_2O_{(l)} \rightarrow Cl_{2\ (aq)} + ClO_3^-{}_{(aq)} + 2OH^-$

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One more example:

Balance the following equation in basic solution:

$I^-_{(aq)} + NO_2^-{}_{(aq)} \rightarrow I_{2\ (s)} + NO_{(g)}$

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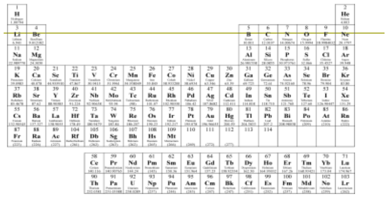
The Periodic Table of the Elements



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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 elementals are 0)

NO_2^- - N is +3

NO – N is +2

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1 – Separate into ½ reactions

$I^-_{(aq)} + NO_2^-{}_{(aq)} \rightarrow I_{2\,(s)} + NO_{(g)}$

Break the full reaction into 2 half-reactions:

Oxidation: $I^-_{(aq)} \rightarrow I_{2\,(s)}$

Reduction: $NO_2^-{}_{(aq)} \rightarrow NO_{(g)}$

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2 – Balance each ½ reaction, ignoring O and H

Oxidation: $I^-_{(aq)} \rightarrow I_{2\,(s)}$

Reduction: $NO_2^-{}_{(aq)} \rightarrow NO_{(g)}$

Just want same number of atoms on each side.

Oxidation: $2\,I^-_{(aq)} \rightarrow I_{2\,(s)}$

Reduction: $NO_2^-{}_{(aq)} \rightarrow NO_{(g)}$

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3 – Balance the oxygen by adding water

Oxidation: $2\,I^-_{(aq)} \rightarrow I_{2\,(s)}$

Reduction: $NO_2^-{}_{(aq)} \rightarrow NO_{(g)}$

Oxidation: $2\,I^-_{(aq)} \rightarrow I_{2\,(s)}$

Reduction: $NO_2^-{}_{(aq)} \rightarrow NO_{(g)} + H_2O_{(l)}$

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4 – Balance the hydrogen by adding H^+

Oxidation: $2\,I^-_{(aq)} \rightarrow I_{2\,(s)}$

Reduction: $NO_2^-{}_{(aq)} \rightarrow NO_{(g)} + H_2O_{(l)}$

Ox: $2\,I^-_{(aq)} \rightarrow I_{2\,(s)}$

Red: $NO_2^-{}_{(aq)} + 2\,H^+_{(aq)} \rightarrow NO_{(g)} + H_2O_{(l)}$

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5 – Balance the charges by adding electrons

Ox: $2\ I^-_{(aq)} \rightarrow I_{2\ (s)}$

Red: $NO_2^-{}_{(aq)} + 2\ H^+_{(aq)} \rightarrow NO_{(g)} + H_2O_{(l)}$

Ox: $2\ I^-_{(aq)} \rightarrow I_{2\ (s)} + 2\ e^-$

Red: $NO_2^-{}_{(aq)} + 2\ H^+_{(aq)} + 1\ e^- \rightarrow NO_{(g)} + H_2O_{(l)}$

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6 – Combine the half-reaction, eliminating any electrons

Ox: $2\ I^-_{(aq)} \rightarrow I_{2\ (s)} + 2\ e^-$

2 electrons

Red: $NO_2^-{}_{(aq)} + 2\ H^+_{(aq)} + 1\ e^- \rightarrow NO_{(g)} + H_2O_{(l)}$

1 electrons

Ox + 2*Red

Ox: $2\ I^-_{(aq)} \rightarrow I_{2\ (s)} + 2\ e^-$

Red: $2*\ (NO_2^-{}_{(aq)} + 2\ H^+_{(aq)} + 1\ e^- \rightarrow NO_{(g)} + H_2O_{(l)})$

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6 – Combine the half-reaction, eliminating any electrons

Ox: $2\ I^-_{(aq)} \rightarrow I_{2\ (s)} + 2\ e^-$

Red: $2*(NO_2^-{}_{(aq)} + 2\ H^+_{(aq)} + 1\ e^- \rightarrow NO_{(g)} + H_2O_{(l)})$

Ox: $2\ I^-_{(aq)} \rightarrow I_{2\ (s)} + 2\ e^-$

Red: $2\ NO_2^-{}_{(aq)} + 4\ H^+_{(aq)} + 2\ e^- \rightarrow 2\ NO_{(g)} + 2\ H_2O_{(l)}$

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6 – Combine the half-reaction, eliminating any electrons

Ox: $2\ I^-_{(aq)} \rightarrow I_{2\,(s)} + 2\ e^-$

Red: $2\ NO_2^-{}_{(aq)} + 4\ H^+_{(aq)} + 2\ e^- \rightarrow 2\ NO_{(g)} + 2\ H_2O_{(l)}$

$2\ I^-_{(aq)} + 2\ NO_2^-{}_{(aq)} + 4\ H^+_{(aq)} + \cancel{2\ e^-} \rightarrow I_{2\,(s)} + \cancel{2\ e^-} + 2\ NO_{(g)} + 2\ H_2O_{(l)}$

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

$2\ I^-_{(aq)} + 2\ NO_2^-{}_{(aq)} + 4\ H^+_{(aq)} \rightarrow I_{2\,(s)} + 2\ NO_{(g)} + 2\ H_2O_{(l)}$

$2\ I^-_{(aq)} + 2\ NO_2^-{}_{(aq)} + 4\ H^+_{(aq)} + 4\ OH^- \rightarrow I_{2\,(s)} + 2\ NO_{(g)} + 2\ H_2O_{(l)} + 4\ OH^-$

Why 4 OH-? Because I need to neutralize 4 H^+ which gives me…4 H_2O!!

$2\ I^-_{(aq)} + 2\ NO_2^-{}_{(aq)} + 4\ H_2O \rightarrow I_{2\,(s)} + 2\ NO_{(g)} + 2\ H_2O_{(l)} + 4\ OH^-$

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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:

$2\ I^-_{(aq)} + 2\ NO_2^-{}_{(aq)} + \overset{2}{\cancel{4}}\ H_2O \rightarrow I_{2\,(s)} + 2\ NO_{(g)} + \cancel{2\ H_2O_{(l)}} + 4\ OH^-$

$2\ I^-_{(aq)} + 2\ NO_2^-{}_{(aq)} + 2\ H_2O \rightarrow I_{2\,(s)} + 2\ NO_{(g)} + 4\ OH^-$

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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 H_2O
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:

$O_{2(g)} \rightarrow H_2O_{(l)}$

$H_2SO_{3\,(aq)} \rightarrow SO_4^{2-}{}_{(aq)}$

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$O_{2(g)} \rightarrow H_2O_{(l)} + H_2O$

$H_2SO_{3\,(aq)} + H_2O \rightarrow SO_4^{2-}{}_{(aq)}$

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$O_{2(g)} + 4H^+ \rightarrow H_2O_{(l)} + H_2O$

$H_2SO_{3\,(aq)} + H_2O \rightarrow SO_4{}^{2-}{}_{(aq)} + 4\,H^+$

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$O_{2(g)} + 4H^+ + 4\,e^- \rightarrow H_2O_{(l)} + H_2O$

$H_2SO_{3\,(aq)} + H_2O \rightarrow SO_4^{2-} + 4\,H^+ + 2e^-$

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$O_{2(g)} + 4H^+ + 4\,e^- \rightarrow H_2O_{(l)} + H_2O$

$2x[H_2SO_3 + H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-]$
$2H_2SO_3 + 2\,H_2O \rightarrow 2SO_4^{2-} + 8H^+ + 4e^-$

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$O_{2(g)} + 4H^+ + 4e^- \rightarrow H_2O_{(l)} + H_2O$

$\underline{2H_2SO_3 + 2H_2O \rightarrow 2SO_4^{2-} + 8H^+ + 4e^-}$

$O_{2(g)} + 4H^+ + 4e^- + 2H_2SO_3 + 2H_2O$

$\rightarrow H_2O_{(l)} + H_2O + 2SO_4^{2-} + 8H^+ + 4e^-$

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$O_{2(g)} + 4H^+ + 4e^- \rightarrow H_2O_{(l)} + H_2O$

$\underline{2H_2SO_3 + 2H_2O \rightarrow 2SO_4^{2-} + 8H^+ + 4e^-}$

$O_{2(g)} + 2H_2SO_3 \rightarrow 2SO_4^{2-} + 4H^+$

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