

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 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 Fe²⁺ 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. Fe²⁺ has 1 more electron than Fe³⁺

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What does this reaction look like?



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



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

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

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

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

Reduction half-reaction: $\text{Mn}^{5+} + 3 \text{e}^- \rightarrow \text{Mn}^{2+}$

Full reaction: $3 \text{Fe}^{2+} + \text{Mn}^{5+} \rightarrow 3 \text{Fe}^{3+} + \text{Mn}^{2+}$

WTFDYG??????????????

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

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

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

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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?

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

CuO – copper (II) oxide
Cu₂O – copper (I) oxide

FeO – iron (II) oxide
Fe₂O₃ – 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?

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

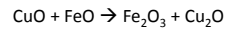
CuO – copper (II) oxide
Cu₂O – copper (I) oxide

FeO – iron (II) oxide
Fe₂O₃ – 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?

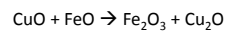


Balancing redox reactions is similar to regular equations BUT it also requires that you balance the charges as well.

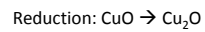
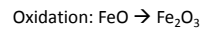
Fortunately, there is a relatively easy system that ALWAYS works! Just follow the 7-ish easy steps!

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1 – Separate into $\frac{1}{2}$ reactions



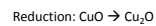
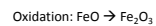
Break the full reaction into 2 half-reactions:



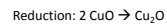
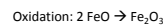
We treat them separately from now on.

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2 – Balance each $\frac{1}{2}$ reaction, ignoring O and H



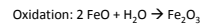
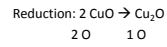
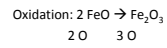
Just want same number of atoms on each side.



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3 – Balance the oxygen by adding water

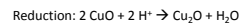
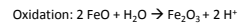
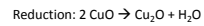
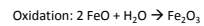
This is more logical than it seems since most electrochemistry occurs in aqueous media.



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4 – Balance the hydrogen by adding H⁺

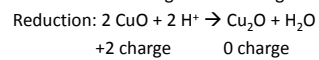
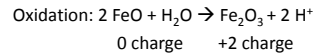
This is also more logical than it seems, since aqueous solutions (as we've seen) are generally either acidic or basic.



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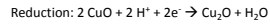
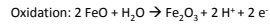
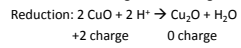
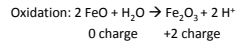
The atoms are balanced

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



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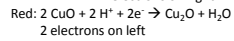
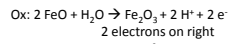
5 – Balance the charges by adding electrons



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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)

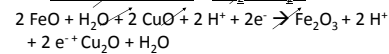
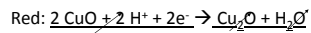
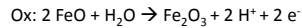


I just add them together as is. If there were a different number of electrons, I'd need to multiply the reactions by whatever factors make them the same.

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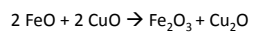
6 – Combine the half-reaction, eliminating any electrons and canceling common components

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



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

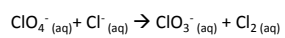


Not a factor here!

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New example:

Balance the following equation in basic solution:

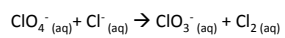


We just need to apply our 7-ish steps.

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New example:

Balance the following equation in basic solution:



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1 – Separate into $\frac{1}{2}$ reactions

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

What's changing oxidation state?
Cl⁻ - oxidation state is -1
Cl₂ - oxidation state is 0 (all elementals are 0)
ClO₄⁻ WTFITOS?
ClO₃⁻ WTFITOS?

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1 – Separate into $\frac{1}{2}$ reactions

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

What's changing oxidation state?
Cl⁻ - oxidation state is -1
Cl₂ - oxidation state is 0 (all elementals are 0)
ClO₄⁻ - Cl is +7 (O is -2, ion is -1 overall)
ClO₃⁻ - Cl is +5 (O is -2, ion is -1 overall)

Slide 27

1 – Separate into $\frac{1}{2}$ reactions

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

Break the full reaction into 2 half-reactions:

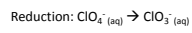
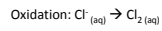
Oxidation: $\text{Cl}^- (\text{aq}) \rightarrow \text{Cl}_2 (\text{aq})$

Reduction: $\text{ClO}_4^- (\text{aq}) \rightarrow \text{ClO}_3^- (\text{aq})$

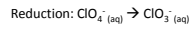
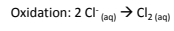
We treat them separately from now on.

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2 – Balance each $\frac{1}{2}$ reaction, ignoring O and H

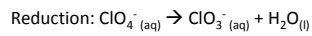
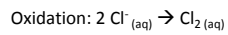
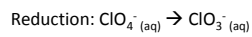
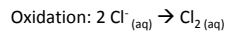


Just want same number of atoms on each side.



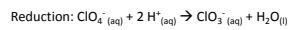
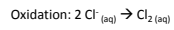
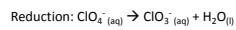
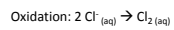
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3 – Balance the oxygen by adding water

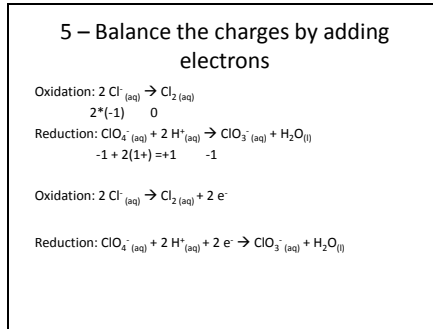


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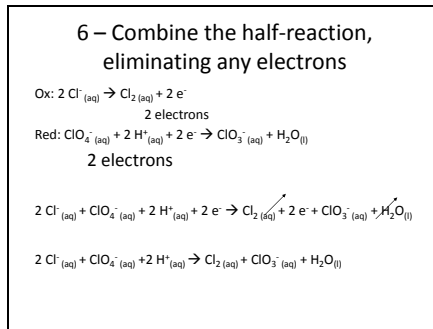
4 – Balance the hydrogen by adding H^+



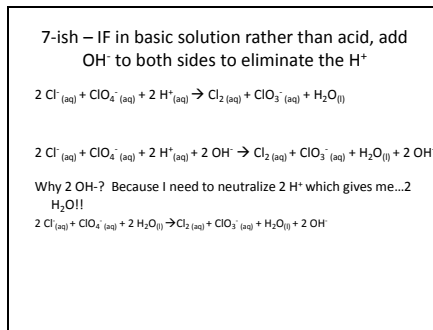
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Slide 33



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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:

$$2 \text{Cl}^-_{(aq)} + \text{ClO}_2^-_{(aq)} + 2 \text{H}_2\text{O}_{(l)} \rightarrow \text{Cl}_2_{(aq)} + \text{ClO}_3^-_{(aq)} + \text{H}_2\text{O}_{(l)} + 2 \text{OH}^-$$
$$2 \text{Cl}^-_{(aq)} + \text{ClO}_2^-_{(aq)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{Cl}_2_{(aq)} + \text{ClO}_3^-_{(aq)} + 2 \text{OH}^-$$

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One more example:

Balance the following equation in basic solution:

$$\text{I}^-_{(aq)} + \text{NO}_2^-_{(aq)} \rightarrow \text{I}_2_{(s)} + \text{NO}_{(g)}$$

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1 – Separate into ½ reactions

$$\text{I}^-_{(aq)} + \text{NO}_2^-_{(aq)} \rightarrow \text{I}_2_{(s)} + \text{NO}_{(g)}$$

What's changing oxidation state?

I⁻ - oxidation state is -1

I₂ - oxidation state is 0 (all elementals are 0)

NO₂⁻ - N is +3

NO - N is +2

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1 – Separate into $\frac{1}{2}$ reactions

$$\text{I}^-_{(\text{aq})} + \text{NO}_2^-_{(\text{aq})} \rightarrow \text{I}_2(\text{s}) + \text{NO}_{(\text{g})}$$

Break the full reaction into 2 half-reactions:

Oxidation: $\text{I}^-_{(\text{aq})} \rightarrow \text{I}_2(\text{s})$

Reduction: $\text{NO}_2^-_{(\text{aq})} \rightarrow \text{NO}_{(\text{g})}$

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2 – Balance each $\frac{1}{2}$ reaction, ignoring O and H

Oxidation: $\text{I}^-_{(\text{aq})} \rightarrow \text{I}_2(\text{s})$

Reduction: $\text{NO}_2^-_{(\text{aq})} \rightarrow \text{NO}_{(\text{g})}$

Just want same number of atoms on each side.

Oxidation: $2 \text{I}^-_{(\text{aq})} \rightarrow \text{I}_2(\text{s})$

Reduction: $\text{NO}_2^-_{(\text{aq})} \rightarrow \text{NO}_{(\text{g})}$

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3 – Balance the oxygen by adding water

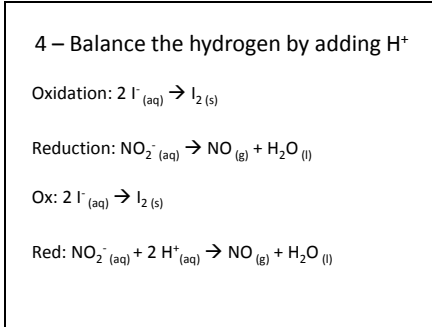
Oxidation: $2 \text{I}^-_{(\text{aq})} \rightarrow \text{I}_2(\text{s})$

Reduction: $\text{NO}_2^-_{(\text{aq})} \rightarrow \text{NO}_{(\text{g})}$

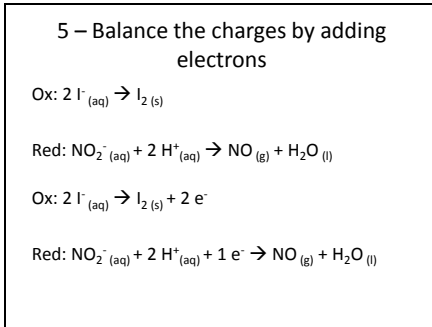
Oxidation: $2 \text{I}^-_{(\text{aq})} \rightarrow \text{I}_2(\text{s})$

Reduction: $\text{NO}_2^-_{(\text{aq})} \rightarrow \text{NO}_{(\text{g})} + \text{H}_2\text{O}_{(\text{l})}$

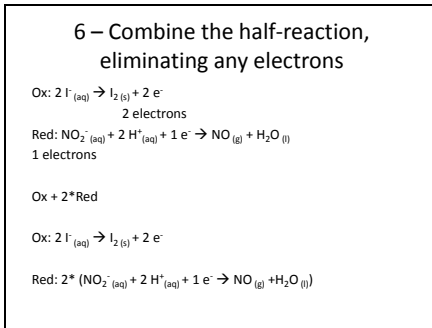
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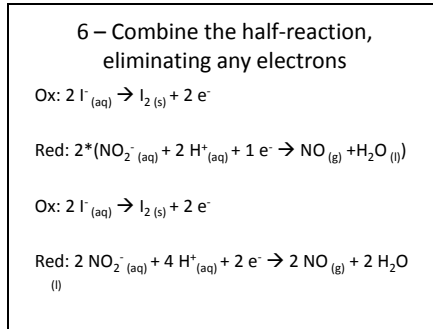
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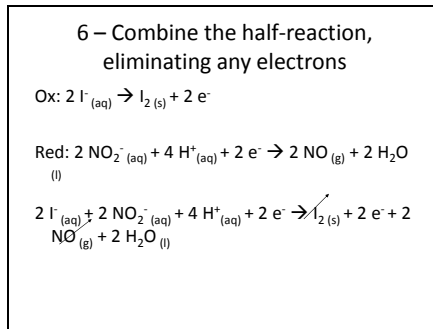
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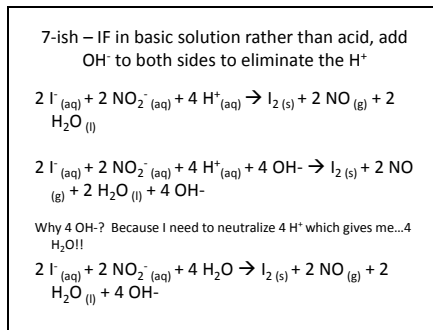
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Slide 45



Slide 46

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

Cleaning up a little bit:

