




Oxidation &
Reduction
The Basics

Presented by
Amelia McCutcheon

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What is Redox?


- Redox is the study of chemical reactions that undergo oxidation and reduction.
- Oxidation cannot occur without reduction, and reduction cannot occur without oxidation.
- This all essentially means that elements lose or gain electrons to change their oxidation state (or oxidation number).



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OIL RIG

- **O**xidation is the **L**oss of electrons
- **R**eduction is the **G**ain of electrons

$$\text{Fe(s)} \rightarrow \text{Fe}^{2+}_{(\text{aq})} + 2\text{e}^{-}$$
$$\text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^{-} \rightarrow \text{Cu}_{(\text{s})}$$


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Addition of half equations


$0 \quad +2$
 $\text{Fe}_{(s)} \rightarrow \text{Fe}^{2+}_{(aq)} + 2e^-$
 Iron is undergoing oxidation – increase in oxidation state

$+2 \quad 0$
 $\text{Cu}^{2+}_{(aq)} + 2e^- \rightarrow \text{Cu}_{(s)}$
 Copper is undergoing reduction – reduction of oxidation state.

YOU MUST BALANCE ELECTRONS

Overall equation

$0 \quad +2 \quad +2 \quad 0$
 $\text{Fe}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Fe}^{2+}_{(aq)} + \text{Cu}_{(s)}$



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Reading the electrochemical series: Electrochemical cells


2. The electrochemical series

E° in volt

$\text{F}_2(\text{g}) + 2e^- \rightleftharpoons 2\text{F}^-(\text{aq})$	+2.87
$\text{Au}^+(\text{aq}) + e^- \rightleftharpoons \text{Au}(\text{s})$	+1.8
$\text{Cl}_2(\text{g}) + 2e^- \rightleftharpoons 2\text{Cl}^-(\text{aq})$	+1.6
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4e^- \rightleftharpoons 2\text{H}_2\text{O}(\text{l})$	+1.3
$\text{Br}_2(\text{l}) + 2e^- \rightleftharpoons 2\text{Br}^-(\text{aq})$	+1.0
$\text{Ag}^+(\text{aq}) + e^- \rightleftharpoons \text{Ag}(\text{s})$	+0.9
$\text{Fe}^{3+}(\text{aq}) + e^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.7
$\text{I}_2(\text{s}) + 2e^- \rightleftharpoons 2\text{I}^-(\text{aq})$	+0.54

Reduction (cathode) ↑

Oxidation (anode) ↓



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Addition of half equations: Ag/Cu cell


$\text{Ag}^+_{(aq)} + e^- \rightarrow \text{Ag}_{(s)} \quad E^\circ = +0.80 \text{ V}$
 $\text{Cu}^{2+}_{(aq)} + 2e^- \rightarrow \text{Cu}_{(s)} \quad E^\circ = +0.34 \text{ V}$

Therefore

$\text{Ag}^+_{(aq)} + e^- \rightarrow \text{Ag}_{(s)} \quad \times 2$
 $\text{Cu}_{(s)} \rightarrow \text{Cu}^{2+}_{(aq)} + 2e^-$

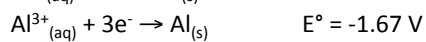
BALANCE ELECTRONS

$\text{Cu}_{(s)} + 2\text{Ag}^+_{(aq)} \rightarrow \text{Cu}^{2+}_{(aq)} + 2\text{Ag}_{(s)}$



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Addition of half equations: Ag/Al cell



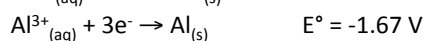
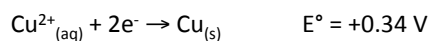
Therefore



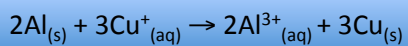
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Addition of half equations: Cu/Al cell



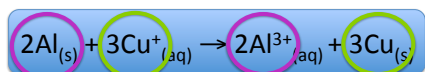
Therefore



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Conjugate Redox Pairs



Conjugate
Redox Pair

Conjugate
Redox Pair



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Calculating oxidation numbers

1. Any pure element – solid liquid or gas – will always have an oxidation number of 0.
2. Any ion, whether aqueous (alone) or part of an ionic compound, will have an oxidation number equal to that of its charge.
3. Hydrogen has an oxidation number of +1, except in metal hydrides, where the oxidation number is -1.
4. Oxygen has an oxidation number of -2, except in peroxides, where the oxidation number is -1.
5. All other oxidation number can be calculated



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Oxidation Numbers

Determine the oxidation numbers of all elements in these compounds:

- | | | |
|---------------------|-----------------------|---------------------------------------|
| 1) Mg | 17) CO ₂ | 33) Fe(NO ₃) ₃ |
| 2) Na ⁺ | 18) O ₃ | 34) H ₃ PO ₄ |
| 3) O ²⁻ | 19) N ₂ | 35) NaH |
| 4) Cr ³⁺ | 20) NO ₂ | 36) CaO |
| 5) Au | 21) XeF ₄ | 37) MnO ₂ |
| 6) Ca | 22) ZnBr ₂ | 38) KMnO ₄ |

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Balancing redox half equations in acidic solution

K Balance **K**ey elements (i.e. all elements other than O and H)

O Balance **O** with H₂O_(l) molecules

H Balance **H** with H⁺_(aq) ions

E Balance charge by adding **E**lectrons



S Add **S**tates

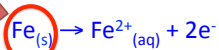
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Oxidants & Reductants

Reductant:
Reducing Agent

– It itself undergoes
oxidation



Oxidant:
Oxidising Agent

– It itself undergoes
reduction



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