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Oxidation & Reduction
The Basics

Presented by
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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).

OIL RIG

- Oxidation is the Loss of electrons
- Reduction is the Gain of electrons

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

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

Addition of half equations

0 +2

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

Iron is undergoing oxidation – increase in oxidation state

+2 0

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

Copper is undergoing reduction – reduction of oxidation state.

YOU MUST BALANCE ELECTRONS

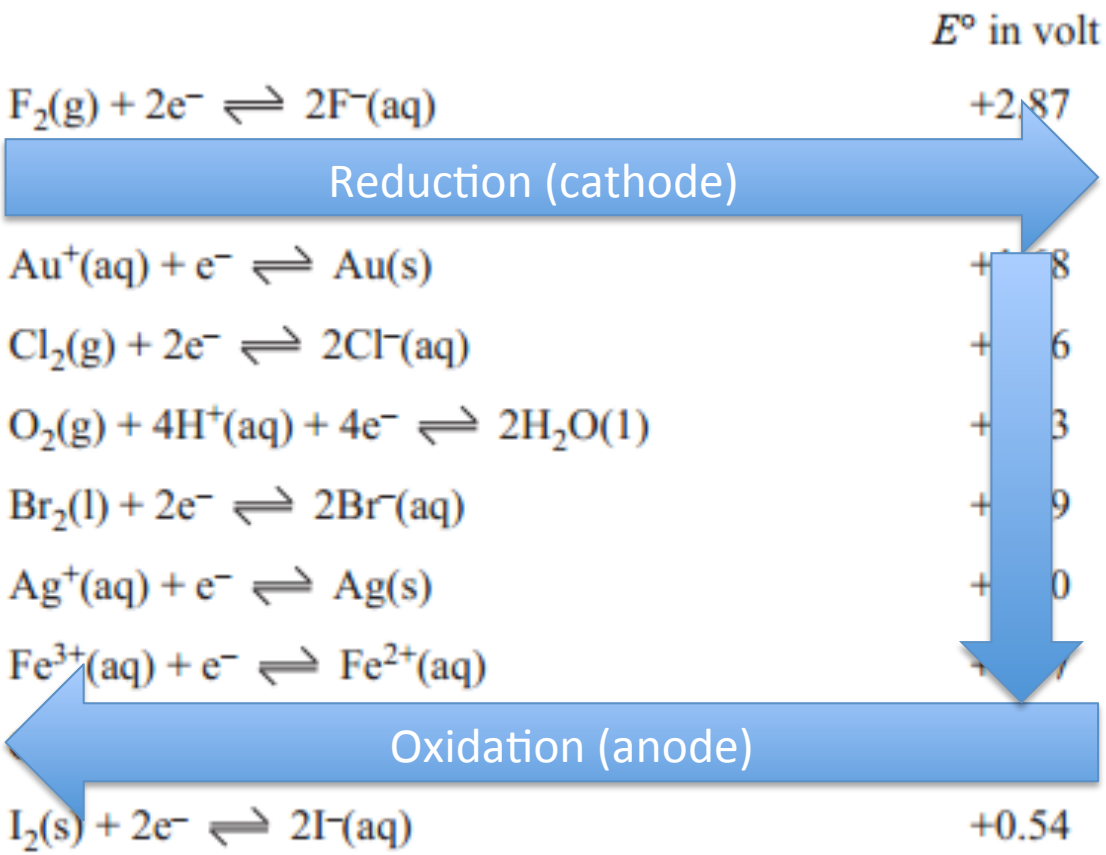
Overall equation

0 +2 +2 0

$Fe(s) + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$

Reading the electrochemical series: Electrochemical cells

2. The electrochemical series



Addition of half equations: Ag/Cu cell

$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ E° = +0.80 V

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ E° = +0.34 V

Therefore

$[Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}]$ x2

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

BALANCE ELECTRONS

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

Addition of half equations: Ag/Al cell

$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ E° = +0.80 V

$Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$ E° = -1.67 V

Therefore

$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ x3

$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$

BALANCE ELECTRONS

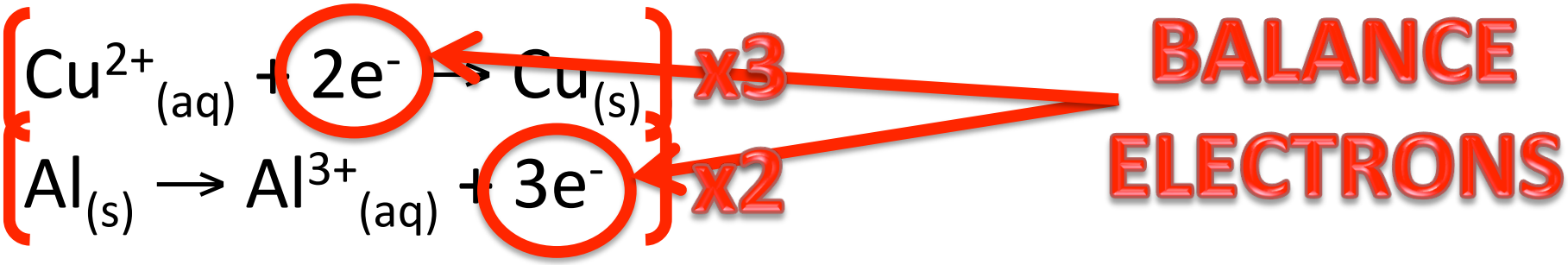
$$Al_{(s)} + 3Ag^+_{(aq)} \rightarrow Al^{3+}_{(aq)} + 3Ag_{(s)}$$

Addition of half equations: Cu/Al cell

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ E° = +0.34 V

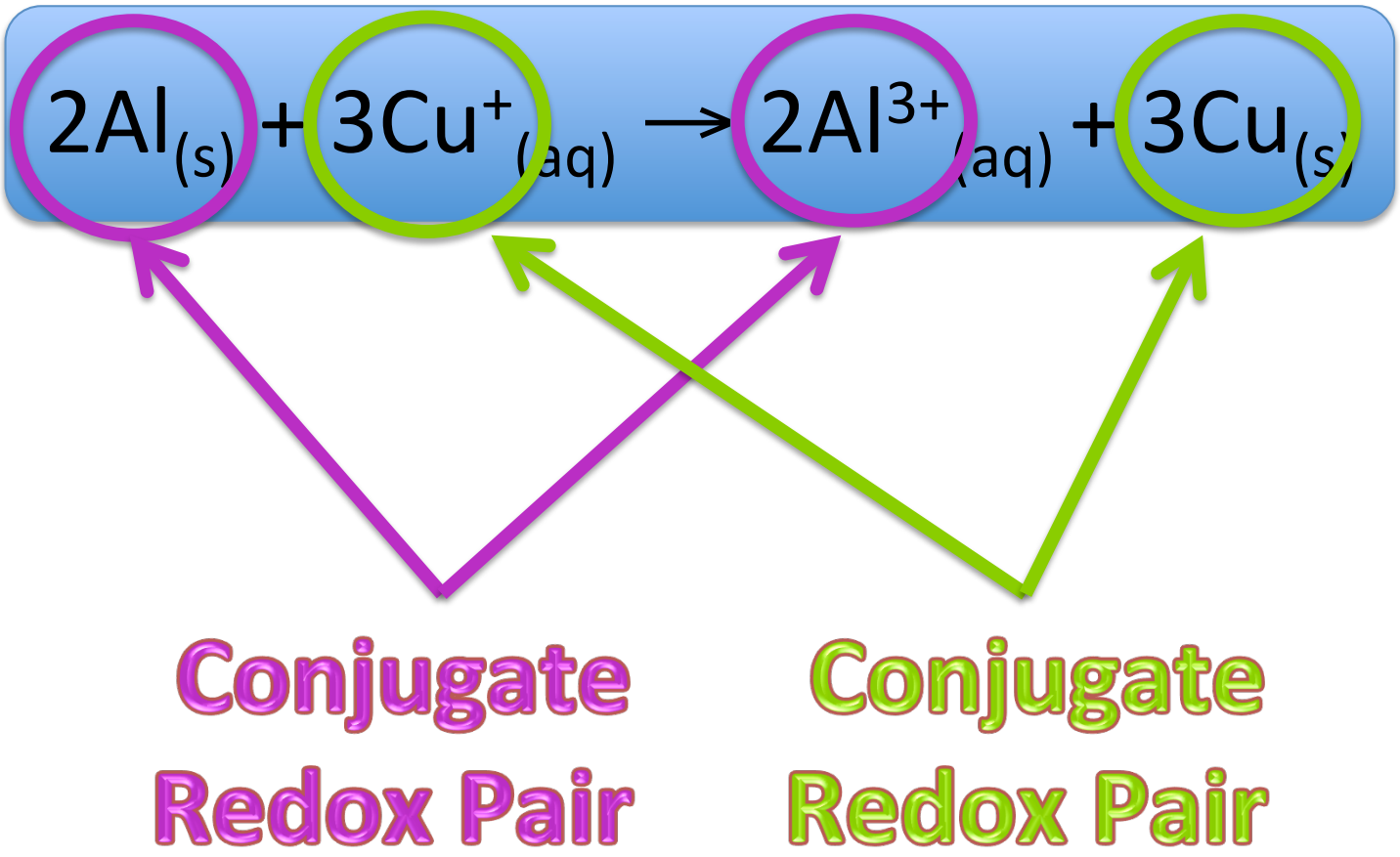
$Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$ E° = -1.67 V

Therefore



$$2Al_{(s)} + 3Cu^{+}_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3Cu_{(s)}$$

Conjugate Redox Pairs



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_3PO_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)	$KMnO_4$

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 $H_2O_{(l)}$ molecules

H Balance **H** with $H^+_{(aq)}$ ions

E Balance charge by adding **E**lectrons

S Add **S**tates

Oxidants & Reductants

- Reductant:
  Reducing Agent
  - It itself undergoes oxidation

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

- Oxidant:
  Oxidising Agent
  - It itself undergoes reduction

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

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