Oxidation-Reduction (Redox) Reactions

☞ A reaction in which one species transfers electrons to another is an ***oxidation-reduction reaction***, also called a ***redox reaction***.

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s) \equiv 2[Fe^{3+}][Cl^-]_3$$

✓ Oxidation is the ***loss*** of electrons by a substance.
✓ Reduction is the ***gain*** of electrons by a substance.

$2(Fe^0 \rightarrow Fe^{3+} + 3e^-)$	Fe^0 "pushes" e's ☞ oxidation
$3(Cl_2 + 2e^- \rightarrow 2Cl^-)$	Cl_2 "pulls" e's ☞ reduction
$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$	redox

☛ ***There is never an oxidation without a reduction, and vice versa!***

Oxidizing Agents and Reducing Agents

✓ An ***oxidizing agent*** (or ***oxidant***) is a substance that causes another substance to be oxidized and is itself reduced.

✓ A ***reducing agent*** (or ***reductant***) is a substance that causes another substance to be reduced and is itself oxidized.

☞ In these terms, all redox reactions take on the general form

$$Ox_1 + Red_2 \rightleftharpoons Red_1 + Ox_2$$

Oxidation Numbers and Redox

▲ When a species is oxidized, one of its atoms goes to a higher (more positive or less negative) oxidation number.

▼ When a species is reduced, one of its atoms goes to a lower (less positive or more negative) oxidation number.

$2\ (Fe^0 \rightarrow Fe^{3+} + 3e^-)$	Fe oxidation number increases ⇒ **oxidation**
$3\ (Cl_2{}^0 + 2e^- \rightarrow 2Cl^-)$	Cl oxidation number decreases ⇒ **reduction**
$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$	redox

Balancing Redox Equations by the Ion-Electron Method

1. **Separate the skeletal equation into two half reactions**. Each half reaction refers to the conversion of a species in either its oxidized or reduced form into a related species in either its reduced or oxidized form. One half reaction will be a reduction and the other will be an oxidation.

2. **Balance each half reaction separately.** Balance atoms on each side of a half reaction by inspection, using H_2O, H^+ (if in acid), or OH^- (if in base) to make the balance in hydrogen and/or oxygen, if needed. Do *not* add any other new species (e.g., O_2, H_2) unless already a part of the skeletal half reaction.

3. **Balance the net charge across each half reaction by adding electrons to the side with the more positive net ionic charge.** If by this process electrons are added on the left side of a half reaction, the half reaction is a reduction. If electrons are added to the right side, the half reaction is an oxidation. (If you add electrons to the same side in both half reactions, something is wrong!)

4. **Multiply both half-reactions by appropriate factors (usually whole numbers), so that the number of electrons is the same in both half reactions and will cancel when the two are added together.**

5. **Add the two multiplied half reactions together to obtain the overall redox equation**.

6. **Check the balance.** No electrons should appear in the overall redox equation. Not only should there be a balance in atoms across the equation, but also the net charge on both sides of the equation should be equal.

Work-Around Technique for Difficult Basic Cases

☹ Balancing H and O in basic redox reactions sometimes can be difficult, because both OH^- and H_2O contain both elements.

☺ A trick to balance troublesome basic cases:

- ✔ Balance any troublesome half-reaction or the entire redox reaction as if it were in acid first.
- ✔ Then add equal numbers of OH^- to both sides of the acid-balanced equation to "neutralize" any H^+ to become H_2O; i.e., $H^+ + OH^- = H_2O$.

Example:

In acid	$6e^- + 6H^+ + IO_3^- \rightarrow I^- + 3H_2O$
Add OH^-	$6OH^- \qquad\qquad\qquad\qquad 6OH^-$
Add	$6e^- + 6H_2O + IO_3^- \rightarrow I^- + 3H_2O + 6OH^-$
Net	$6e^- + 3H_2O + IO_3^- \rightarrow I^- + 6OH^-$