Oxidation States of Atoms in Elements and Compounds

		Oxidation Number	Examples
1	**Atoms in their elemental state**	0	Fe, H_2, O_2
2	**Monatomic ions**	charge	F^-, Na^+, Fe^{3+}
3	**Group 1A**	+1	NaCl, KNO_3
4	**Group 2A**	+2	MgO
5	**Fluorine**	-1	HF, ClF
6	**Hydrogen**	+1	H_2O
7	**Oxygen**	-2	SO_2, $HClO_4$
8	**Group 7A**	-1	HCl
9	**Group 6A**	-2	PbS_2

Balancing Oxidation-Reduction Reactions

1. Assign oxidation numbers.

2. Separate into oxidation and reduction half reactions.

3. Balance each half reaction using the following steps:
   a. Balance all elements except oxygen or hydrogen.
   b. Balance oxygen by adding H_2O.
   c. Balance hydrogen by adding H^+.
   d. Balance charge by adding electrons:
      Electrons go on the RIGHT (product side) for OXIDATION reactions.
      Electrons go on the LEFT (reactant side) for REDUCTION reactions.
   e. In BASIC solution, do this additional step:
      For every H^+, add OH^- to BOTH sides of the reaction.
      Combine H^+ and OH^- into H_2O.
      Cancel out any waters that appear on both sides.
   You should now have a balanced half reaction.

4. Multiply balanced half reactions so an equal number of electrons are consumed and produced.

5. Add together half reactions.

6. Clean up. Combine identical substances and reduce coefficients to the lowest terms.

7. CHECK! Atom and charge must balance.

Hints: Never add O_2, O atoms, or O^{2-} to balance oxygen.
Never add H_2 or H atoms to balance hydrogen.
Be sure to write the correct charges on all the ions.

Table 20.1 Standard Reduction Potentials in Aqueous Solution at 25 °C*

Reduction Half-Reaction		$E°$ (V)
$F_2(g) + 2\ e^-$	$\longrightarrow 2\ F^-(aq)$	+2.87
$H_2O_2(aq) + 2\ H^+(aq) + 2\ e^-$	$\longrightarrow 2\ H_2O(\ell)$	+1.77
$PbO_2(s) + SO_4^{2-}(aq) + 4\ H^+(aq) + 2\ e^-$	$\longrightarrow PbSO_4(s) + 2\ H_2O(\ell)$	+1.685
$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^-$	$\longrightarrow Mn^{2+}(aq) + 4\ H_2O(\ell)$	+1.51
$Au^{3+}(aq) + 3\ e^-$	$\longrightarrow Au(s)$	+1.50
$Cl_2(g) + 2\ e^-$	$\longrightarrow 2\ Cl^-(aq)$	+1.36
$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 6\ e^-$	$\longrightarrow 2\ Cr^{3+}(aq) + 7\ H_2O\ (\ell)$	+1.33
$O_2(g) + 4\ H^+(aq) + 4\ e^-$	$\longrightarrow 2\ H_2O(\ell)$	+1.229
$Br_2(\ell) + 2\ e^-$	$\longrightarrow 2\ Br^-(aq)$	+1.08
$NO_3^-(aq) + 4\ H^+(aq) + 3\ e^-$	$\longrightarrow NO(g) + 2\ H_2O(\ell)$	+0.96
$OCl^-(aq) + H_2O(\ell) + 2\ e^-$	$\longrightarrow Cl^-(aq) + 2\ OH^-(aq)$	+0.89
$Hg^{2+}(aq) + 2\ e^-$	$\longrightarrow Hg(\ell)$	+0.855
$Ag^+(aq) + e^-$	$\longrightarrow Ag(s)$	+0.799
$Hg_2^{2+}(aq) + 2\ e^-$	$\longrightarrow 2\ Hg(\ell)$	+0.789
$Fe^{3+}(aq) + e^-$	$\longrightarrow Fe^{2+}(aq)$	+0.771
$I_2(s) + 2\ e^-$	$\longrightarrow 2\ I^-(aq)$	+0.535
$O_2(g) + 2\ H_2O(\ell) + 4\ e^-$	$\longrightarrow 4\ OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2\ e^-$	$\longrightarrow Cu(s)$	+0.337
$Sn^{4+}(aq) + 2\ e^-$	$\longrightarrow Sn^{2+}(aq)$	+0.15
$2\ H^+(aq) + 2\ e^-$	$\longrightarrow H_2(g)$	0.00
$Sn^{2+}(aq) + 2\ e^-$	$\longrightarrow Sn(s)$	−0.14
$Ni^{2+}(aq) + 2\ e^-$	$\longrightarrow Ni(s)$	−0.25
$V^{3+}(aq) + e^-$	$\longrightarrow V^{2+}(aq)$	−0.255
$PbSO_4(s) + 2\ e^-$	$\longrightarrow Pb(s) + SO_4^{2-}(aq)$	−0.356
$Cd^{2+}(aq) + 2\ e^-$	$\longrightarrow Cd(s)$	−0.40
$Fe^{2+}(aq) + 2\ e^-$	$\longrightarrow Fe(s)$	−0.44
$Zn^{2+}(aq) + 2\ e^-$	$\longrightarrow Zn(s)$	−0.763
$2\ H_2O(\ell) + 2\ e^-$	$\longrightarrow H_2(g) + 2\ OH^-(aq)$	−0.8277
$Al^{3+}(aq) + 3\ e^-$	$\longrightarrow Al(s)$	−1.66
$Mg^{2+}(aq) + 2\ e^-$	$\longrightarrow Mg(s)$	−2.37
$Na^+(aq) + e^-$	$\longrightarrow Na(s)$	−2.714
$K^+(aq) + e^-$	$\longrightarrow K(s)$	−2.925
$Li^+(aq) + e^-$	$\longrightarrow Li(s)$	−3.045

Increasing strength of oxidizing agents ↑

Increasing strength of reducing agents ↓

* In volts (V) versus the standard hydrogen electrode.

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