# Oxidation – Reduction (aka Redox) Reactions

**One or more electrons are transferred in a redox reaction**

Examples: photosynthesis, oxidation of sugars, fats, and proteins, and combustion reactions

To keep track of the electrons transferred in a redox reaction, we assign oxidation states.

**Oxidation State** – arbitrary charge assigned to an atom

| The Oxidation State of… | Summary | Examples |
|---|---|---|
| An atom in an element is zero | Element: 0 | Na(s), $O_2$(g), Hg(l) |
| A monatomic ion is the same as its charge | Monatomic ion: charge of ion | $Na^+$, $Cl^-$ |
| Fluorine is -1 in its compounds | Fluorine: -1 | HF, $PF_3$ |
| Oxygen is -2 in its compounds<br>*Exception: oxygen is -1 in peroxides ($O_2^{-2}$) | Oxygen: -2<br>*Except in peroxides: O is -1 | $H_2O$, $CO_2$<br>*Exception: $H_2O_2$ |
| Hydrogen is +1 in its covalent compounds | Hydrogen: +1 | $H_2O$, HCl, $NH_3$ |
| Hydrogen is -1 in binary metal hydrides | Hydrogen: -1 | NaH, $MgH_2$ |

**Neutral Compounds:** sum of all oxidation states must equal zero

+1 -2

Examples: $H_2O$ is neutral: each H is +1 and O is -2; SUM = 2(+1) + (-2) = 0

NaCl is neutral: Na is +1 and Cl is -1; SUM = (+1) + (-1) = 0

**Ionic species:** sum of all oxidation states must equal charge of the ion

Examples: $NO_3^-$ has charge of -1; each O is -2 and SUM = -1

N + 3(-2) = -1

N = +5

$Cr_2O_7^{-2}$ has charge of -2; each O is -2 and SUM = -2

2(Cr) + 7(-2) = -2

Cr = +6

$2Cr + (-14) = -2$

$Cr = +6$

1. $CaBr_2$

2. $N_2$

3. $Na_2CO_3$

$(CO_3)^{2-}$

4. $ClO_3^-$

$+5 - 6 = -1$

5. $KMnO_4$ (+1, +7, -2)

6. $SF_6$ (+6, -1)

7. $Fe_3O_4$ ($+\frac{8}{3}$, -2)

8. $CH_3F$ (-2, +1, -1)

*LEO says GER

**Oxidation** – oxidation number increases, loss of electrons

$Pb(s) \rightarrow Pb^{2+}(aq) + 2\,e^-$

$Cr^{2+}(aq) \rightarrow Cr^{3+}(aq) + e^-$

**Reduction** – oxidation number decreases, gain of electrons

$F_2(g) + 2\,e^- \rightarrow 2\,F^-(aq)$

$Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$

LEO = loss of $e^-$ is oxidation

GER = gain of $e^-$ is reduction

OIL = oxidation is loss

RIG = reduction is gain

Half-Box Method

## Guidelines for Balancing Redox Reactions

1. Assign oxidation numbers to each element in the reaction

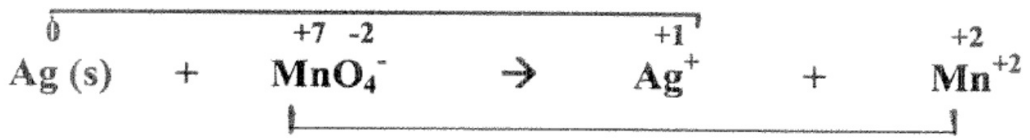
   Review of rules: a. elements = 0 b. ions = charge, c. elements in cpds F = -1, O = -2, H = +1 d. in cpds the oxid #'s of all elements must add up to = 0
   e. in polyions the oxid #'s of all elements must add up to = the charge of the ion

$$\overset{0}{Ag\,(s)} + \overset{+7\ -2}{MnO_4^-} \rightarrow \overset{+1}{Ag^+} + \overset{+2}{Mn^{+2}}$$

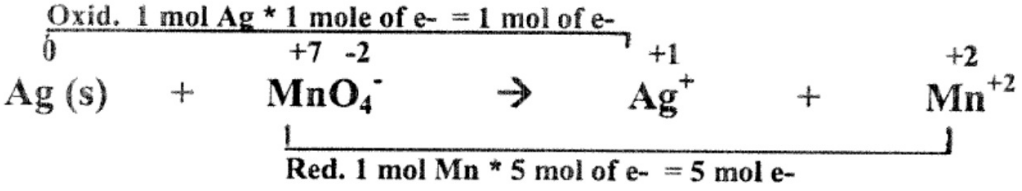
2. Once oxidation numbers are determined, analyze the #'s to see which element is being oxidized and which element is being reduced (remember oxidized means the oxid # goes up and reduced means the oxid # goes down).

***Ag is being oxid. from 0 to +1 and Mn is being red. from +7 to +2***

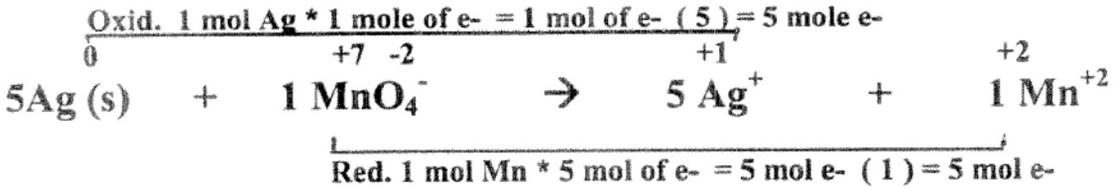
3. Then draw a ½ rectangle box on the top of the reaction from the element being oxidized on the left to its new self on the right. Next draw a ½ rectangle box on the bottom of the reaction from the element being reduced on the left to its new self on the right.



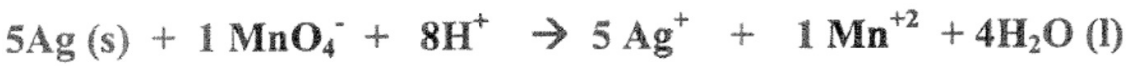

4. Next count the number of things being oxidized on each side as well as the number of electrons transferred (based upon the oxidation # change) and say # moles of element * # moles of electrons transferred = moles of electrons transferred. Repeat this process for the reduction part putting this equation on the top side or bottom side of the appropriate ½ box.




5. The number of electrons transferred must be equal so the two numbers must be multiplied together using ( ) and the number in ( ) must be applied (multiplied) onto the existing coefficients that the appropriate ½ box is pointed to on each side.




6. Finally add the needed number of $H_2O$'s to the side with either no oxygens or less oxygens so that the oxygens are balanced. Then add $H^+$ to the side with no hydrogens or less hydrogens to balance the hydrogens.



7. *If* the reaction happens to take place in ***Basic*** solution instead of acidic solution do this last thing: Add to each side the same number of $OH^-$ that you have of $H^+$ and then combine the $OH^-$ with the $H^+$ that are on the same side to make $H_2O$ on one side and then cancel out the same number of $H_2O$ on each side as a clean up.

$$5Ag\ (s) + MnO_4^- + 8H^+ + 8OH^- \rightarrow 5\ Ag^+ + Mn^{+2} + 4H_2O\ (l) + 8OH^-$$

$$5Ag\ (s) + MnO_4^- + 8H_2O\ (l) \rightarrow 5\ Ag^+ + Mn^{+2} + 4H_2O\ (l) + 8OH^-$$

$$5Ag\ (s) + MnO_4^- + 4H_2O\ (l) \rightarrow 5\ Ag^+ + Mn^{+2} + 8OH^-$$

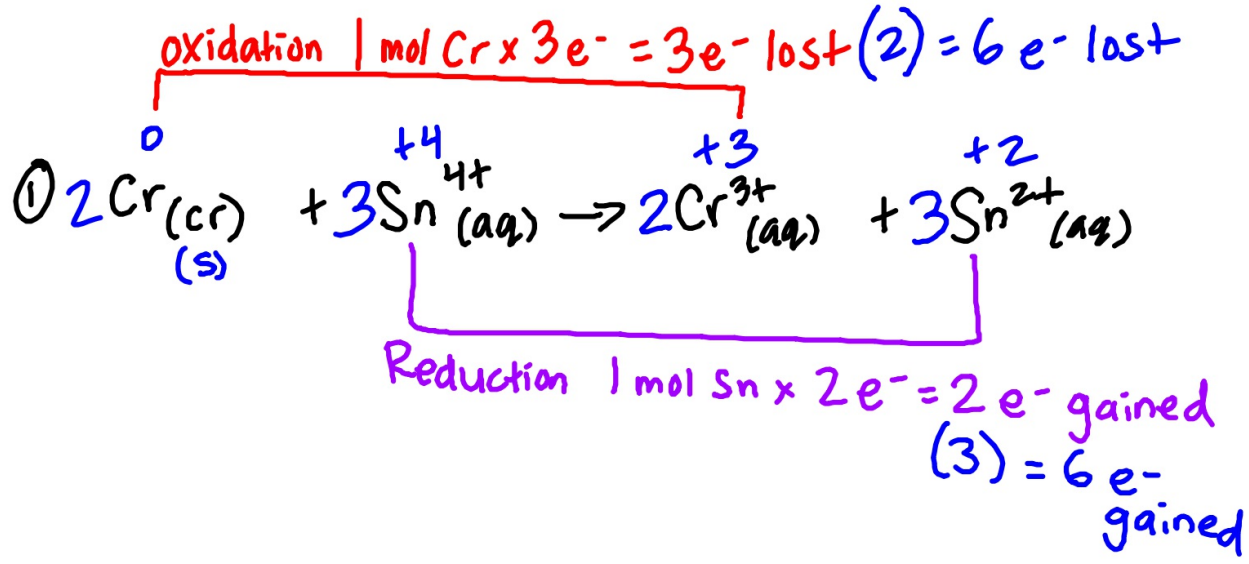
# Balancing Redox Equations

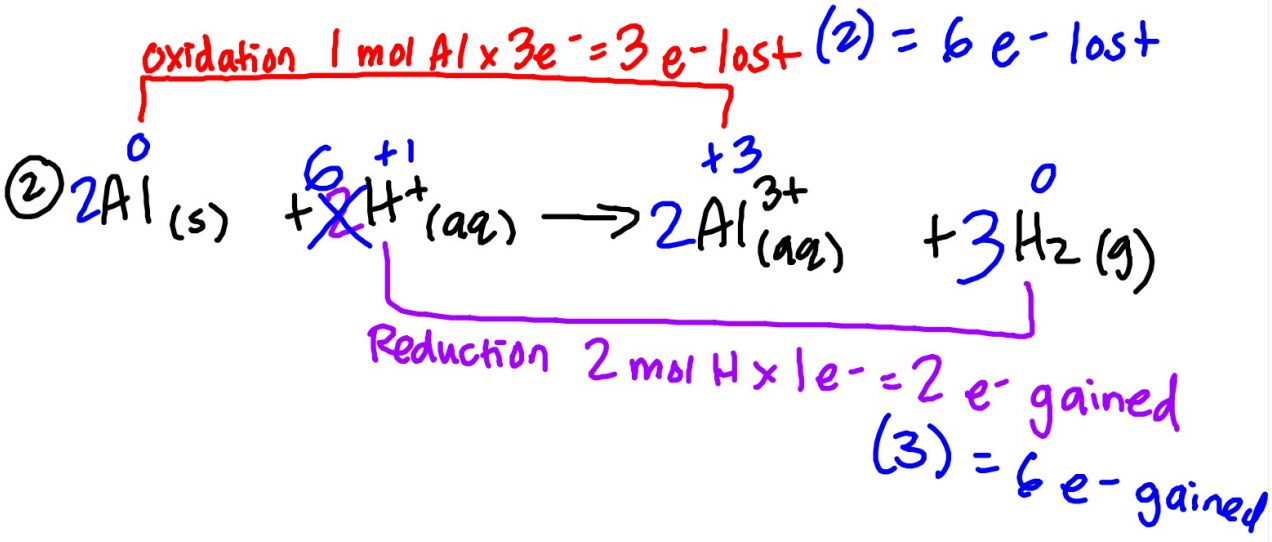
Balance the following REDOX reactions (cr = crystalline solid).

*All reactions take place in an acidic solution unless otherwise indicated.*

1. $Cr\ (cr) + Sn^{4+}\ (aq) \longrightarrow Cr^{3+}\ (aq) + Sn^{2+}\ (aq)$

2. $Al\ (cr) + H^+\ (aq) \longrightarrow Al^{3+}\ (aq) + H_2\ (g)$


oxidation 1 mol Cr x 3e⁻ = 3e⁻ lost (2) = 6 e⁻ lost
① 2Cr(cr) (s) + 3Sn⁴⁺(aq) → 2Cr³⁺(aq) + 3Sn²⁺(aq)
Reduction 1 mol Sn x 2e⁻ = 2 e⁻ gained
(3) = 6 e⁻ gained


oxidation 1 mol Al x 3e⁻ = 3 e⁻ lost (2) = 6 e⁻ lost
② 2Al(s) + 6H⁺(aq) → 2Al³⁺(aq) + 3H₂(g)
Reduction 2 mol H x 1e⁻ = 2 e⁻ gained
(3) = 6 e⁻ gained