- Hebden 4.1 - 4.2

# 1. Oxidation vs. Reduction

## *a) Two Reactions in One!*

i) $Al_{(s)} + Fe_2O_{3\,(s)} \rightarrow Fe_{(l)} + Al_2O_{3\,(s)}$

ii) There are <u>two</u> "half" reactions occurring!

① $Al_{(s)} \rightarrow Al^{+3}$ ($Al_{(s)}$ loses 3 electrons)

② $Fe^{+3} \rightarrow Fe_{(l)}$ ($Fe^{+3}$ gains 3 electrons)

iii) We must balance the half reactions charge using electrons:

① $Al_{(s)} \rightarrow Al^{+3} + 3e^-$ (oxidation)

② $Fe^{+3} + 3e^- \rightarrow Fe_{(l)}$ (reduction)

iv) **Oxidation = Loss of electrons**

"LEO the lion says GER"

**Reduction = Gain of electrons**

v) Oxidation and Reduction half reactions always occur together.

Called a **Redox reaction**.

$3e^- + Al_{(s)} + Fe^{+3} \rightarrow Al^{+3} + Fe_{(l)} + 3e^-$ (cancel e-)

vi) **Oxidizing Agent** = substance doing the oxidation of other species

$Fe^{+3}$ is the oxidizing agent

*(Hint: it's the substance being reduced!)*

**Reducing Agent** = substance doing the reduction of other species

$Al_{(s)}$ is the reducing agent

*(Hint: it's the substance being oxidized!)*

## *b) Example:* $Cu^{+2} + Mg_{(s)} \rightarrow Mg^{+2} + Cu_{(s)}$

i) What are the half reactions?

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

$Mg_{(s)} \rightarrow Mg^{+2} + 2e^-$

ii) What is being oxidized? $Mg_{(s)}$ iii) What is being reduced? $Cu^{+2}$

iv) What is the oxidizing agent? $Cu^{+2}$ v) What is the reducing agent? $Mg_{(s)}$

## *c) Example:* $Ag^+ + Na_{(s)} \rightarrow Ag_{(s)} + Na^+$

i) What are the half reactions?

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

$Na_{(s)} \rightarrow Na^+ + e^-$

ii) What is being oxidized? $Na_{(s)}$ iii) What is being reduced? $Ag^+$

iv) What is the oxidizing agent? $Ag^+$ v) What is the reducing agent? $Na_{(s)}$

## *d) Example:* $PbCl_{2\,(s)} + Fe_{(s)} \rightarrow Pb_{(s)} + FeCl_{2\,(s)}$

i) What are the half reactions?

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

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

ii) What is being oxidized? $Fe_{(s)}$ iii) What is being reduced? $Pb^{+2}$

iv) What is the oxidizing agent? $Pb^{+2}$ v) What is the reducing agent? $Fe_{(s)}$

Do Questions: #1-2 page 192 (Hebden); #1-4 page 621 (Heath)

# 3. Balancing Redox Reactions – Using Oxidation Numbers

## *a) What are "Oxidation Numbers"?*

**Sample: $K_2Cr_2O_7$ K is +1 Cr is +6 O is -2**

- They are not actual charges!
- They are a bookkeeping system used to keep track of an atoms electrons!

### i) Rules for Assigning Oxidation Numbers

1. Oxidation number of an atom as an isolated ion is the charge of the ion.

   Eg: $Cu^{2+}$ has oxidation number of +2

2. Oxidation number of an atom in elemental form is zero.

   Eg. $Cu_{(s)}$ has oxidation number of 0

   Eg. $O_{2\ (g)}$ has oxidation number of 0

   Eg. $S_{8\ (s)}$ has oxidation number of 0

3. Oxidation of *common* atoms in a compound is its combining capacity

   Common atoms include:

   Group 1 (+1); Group 2 (+2); Halogens (-1); Oxygen (-2); Aluminum (+3)

   Eg. Oxidation number of Al in $Al_2O_3$ is +3 (O is –2 in $Al_2O_3$ !)

   Eg. Oxidation number of Na in NaF is +1 (F is –1 in NaF)

4. The sum of the positive and negative charges in a compound equals zero

   Eg. $K_2Cr_2O_7$ K is +1, and O is –2 from rule 3.

   To find oxidation number of Cr:
   $2(+1) + 2(x) + 7(-2) = 0$
   $2 + 2x - 14 = 0$
   $x = +6$

5. The sum of the positive and negative charges in a polyatomic ion equals the charge of the ion.

   Eg. $MnO_4^-$ O is –2 from rule 3

   To find oxidation number of Mn:

   $x + 4(-2) = -1$
   $x - 8 = -1$
   $x = +7$

ii) Example: What are the oxidation numbers for all the atoms in $Al(NO_3)_3$?

$Al = +3, N = +5, O = -2$

iii) Example: What are the oxidation numbers for each atom in $ClO_4^-$? *(Tricky!)*

$Cl = +7, O = -2$ *(oxygen "beats out" the halogen!)*

iv) Example: What are the oxidation numbers for each atom in $KH_2PO_4$?

$K = +1, H = +1, P = +5, O = -2$

v) Example: Use changes in oxidation number to identify which species is oxidized and which is reduced:

$$\overset{0}{Cu}_{(s)} + \overset{+1\ +5\ -2}{2AgNO_3}_{(aq)} \rightarrow \overset{+2\ +5\ -2}{Cu(NO_3)_2}_{(aq)} + \overset{0}{2Ag}_{(s)}$$

Assign oxidation numbers as above. Then look for change in ox. number!

Cu went from 0 to +2. This is a loss of electrons. Oxidation!
Ag+1 went from +1 to 0. This is a gain of electrons. Reduction!

vi) Example: Identify the oxidizing agent and the reducing agent:

$$\overset{+1\ -2}{3H_2S}_{(g)} + \overset{+1\ +5\ -2}{2HNO_3}_{(aq)} \rightarrow \overset{0}{3S}_{(s)} + \overset{+2\ -2}{2NO}_{(g)} + \overset{+1\ -2}{4H_2O}_{(l)}$$

S went from –2 to 0. This is a loss of electrons. Oxidation!
Thus $H_2S$ is oxidized or is the reducing agent!

N went from +5 to +2. This is gain of electrons. Reduction!
Thus, $HNO_3$ is reduced or is the oxidizing agent!

Do Questions: #3-6 page 194-195; #20-23 page 204