REDOX Reactions

Reading:	Ch 4 section 9	Homework:	Chapter 4: 87*, 89, 91*

* = 'important' homework question

Background

'REDOX' reactions are chemical processes in which REDuction and OXidation simultaneously occur

Oxidation Is Loss of electrons. An element or compound that *loses* electron(s) during a chemical process is said to be OXIDIZED

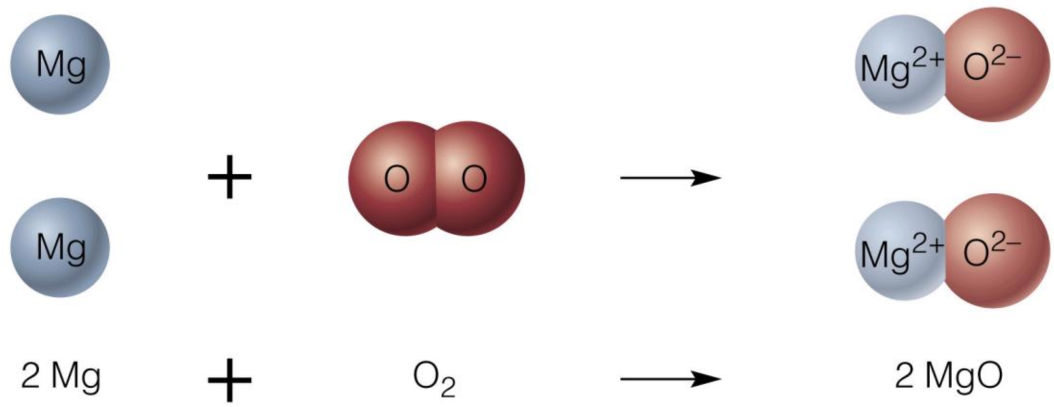
Reduction Is Gain of electrons. An element or compound that *gains* electron(s) during a chemical process is said to be REDUCED

TRICK: Remembering the difference between oxidation and reduction is easy, just remember….

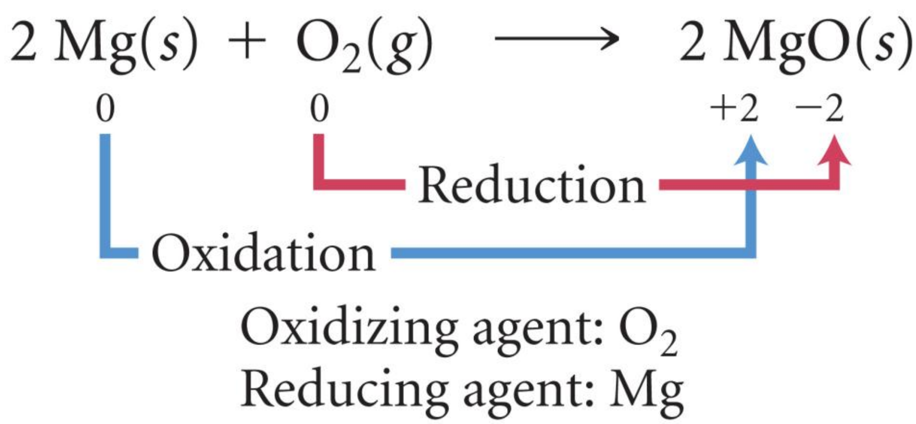
OIL RIG of electrons

Oxidation Is Loss, Reduction Is Gain *of electrons*

Overview Example of a Simple REDOX (combustion) reaction:



Discussion: Which chemical species has lost electrons during this process (i.e. been oxidized)? Which has gained electrons (been reduced)? How can you figure this out?



Oxidized SPECIES =

Oxidizing AGENT =

Reduced SPECIES =

Reducing AGENT =

Notes

Oxidation Numbers

<u>Oxidation Number or Oxidation State</u> – the 'charge' on an atom in a substance as if it were a monatomic ion

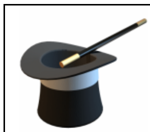
A *change* in oxidation state during a chemical process indicates that a specie has either been *oxidized* (number goes up), or *reduced* (number goes down). Recall the previous MgO example.

<u>Rules for assigning oxidation numbers</u>

1. For materials that form atomic ions, the oxidation state is the same as the 'regular' ionic charge

<u>Task</u>: State the oxidation state of the following:

Na in NaCl		Cl in $AlCl_3$	
Mg in $MgCl_2$		Fe in Fe_2O_3	



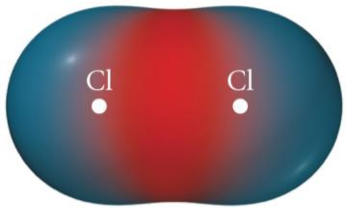
Since you know the charge of a great many atomic ions, you also know their oxidation states, i.e.:

Group I = I (Li^+, Na^+…) **Group VII = -I (F^-, Cl^-…)**
Group II = II (Mg^{2+}, Ca^{2+}…) **Group VI = -II (O^{2-}, S^{2-}…)**

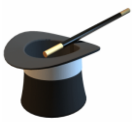
<u>Note</u>: Oxidation states / numbers are expressed with Roman Numerals (this differentiates them from 'pure' ionic charges)

2. For ANY elemental atom, its oxidation state is ZERO. Why?

E.g. Elemental chlorine, Cl_2 (Cl — Cl)



Hint: Recall the trend in Electronegativity values (see appendix)



Any atom bonded to other *identical* atom(s) must have an oxidation state (oxidation number) of ZERO

ALL ELEMENTS must by definition posses zero oxidation states

Examples: Any diatomic element (O_2, F_2), any metallic element (Pb(s), Al(s)) etc.

3. All other atoms' oxidation states must be determined mathematically using the '<u>Sum of Oxidation States</u>' Rule:

<u>For molecules</u>: The sum of the molecule's component atoms individual oxidation numbers = ZERO

Example: Nitric acid, HNO_3

<u>For polyatomic ions (‘charged molecules’)</u>: The sum of the polyatomic ion’s component atoms individual oxidation numbers = overall ionic charge

<u>Example</u>: The nitrate ion, NO_3^-

<u>Exceptions</u>:

Oxygen always has a -II oxidation state, except when bonded to either fluorine or itself. Why? <u>Hint</u>: Think of the periodic trend in electronegativity (appendix).

<u>Examples</u>:

<u>H_2O_2</u>

<u>OF_2</u>

F always has a –I oxidation state, except when bonded to itself. Other Halogens (Cl, Br, I) are also –I, except when bonded to F or O. Why?

Examples:

ClO_2 ClO_3^-

More Examples: Calculate the oxidation state of:

S in SO_3 Xe in XeF_6

S in SO_4^{2-} Cr in CrO_4^-

Na in NaH (Hint: think electronegativity) N in Mg_3N_2

Understanding REDOX reactions

> **Since ‘*battery*’ and *single replacement* reactions are examples of REDOX processes, oxidation numbers can be assigned.**
>
> **Comparing oxidation numbers of reactants and products shows which species have been oxidized and which reduced.**

Example: Zn/Fe^{2+} ‘battery’ style REDOX reaction

$$\begin{array}{ccccccc} Zn(s) & + & Fe^{2+}(aq) & \longrightarrow & Zn^{2+}(aq) & + & Fe(s) \\ 0 & & +2 & & +2 & & 0 \end{array}$$

Reduction

Oxidation

Oxidizing agent: Fe^{2+}
Reducing agent: Zn

Recall: For single replacement (‘prom’) reactions, the more reactive specie (A) undergoes *oxidation* when replacing the less reactive one (B), which is *reduced*.

$$A + BX \rightarrow AX + B$$

Example:

$$\begin{array}{lccccccc} & Mg(s) & + & 2\,H_2O(l) & \longrightarrow & Mg(OH)_2(aq) & + & H_2(g) \\ \text{Oxidation states:} & 0 & & +1\ -2 & & +2\ \ -2\ +1 & & 0 \end{array}$$

Reduction

Oxidation

Tasks:

Assign oxidation numbers to the reactants and products in the following examples. State which specie is oxidized and which is reduced in each case.

$Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Cu\ (s) + Zn^{2+}$

$Ca\ (s) + 2\ HCl\ (aq) \rightarrow CaCl_2 + H_2\ (g)$

$Fe\ (s) + Ni(NO_3)_2\ (aq) \rightarrow Ni\ (s) + Fe(NO_3)_2\ (aq)$

$SO_3\ (g) + H_2O\ (l) \rightarrow H_2SO_4(aq)$

Appendix: Electronegativity values