Chemistry 6A F2007

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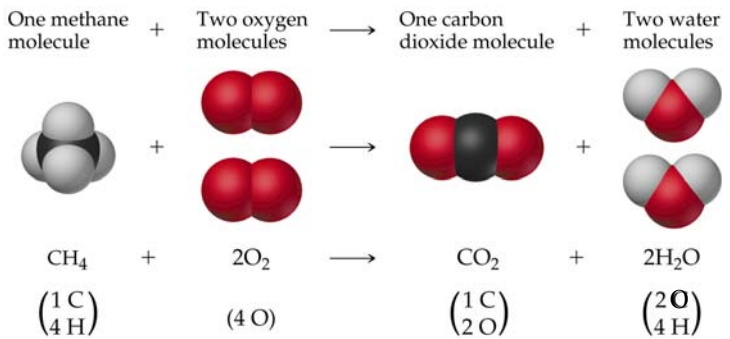
Law of conservation of mass: Matter cannot be lost in any chemical reactions.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$



When a compound contains both carbon, hydrogen & sometimes oxygen, the products are $CO_2(g)$ & $H_2O(l)$

Example:

Methane, $CH_4(g)$ combusts to form carbon dioxide and water

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$$

reacts with oxygen

The products of combustion of a hydrocarbon are always $CO_2(g)$ and $H_2O(l)$

Balancing Combustion Reactions: Example

$$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$$

2 C's & 6 H's | 2 O's | 1 C & 2 O's | 2 H's & 1 O

balance last

balance H first

$$\underline{2}\,C_2H_6 + O_2 \rightarrow CO_2 + \underline{6}\,H_2O$$

balance C next

$$2C_2H_6 + O_2 \rightarrow \underline{4}\,CO_2 + 6H_2O$$

balance O

$$2C_2H_6 + \underline{7}\,O_2 \rightarrow 4CO_2 + 6H_2O$$

4 C's 12 H's 14 O's | *4 C's 12 H's 14 O's*

Chemical Equations:

In a chemical equation one uses chemical symbols to describe a chemical process.

Example: Upon heating, calcium carbonate decomposes to form calcium oxide and carbon dioxide.

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

reactant — *heat supplied* — products

Chemical Equations:

Note that when I wrote the chemical formulas, I also indicated the ***physical state*** of the compound.

One must indicate the state of a compound or element by:

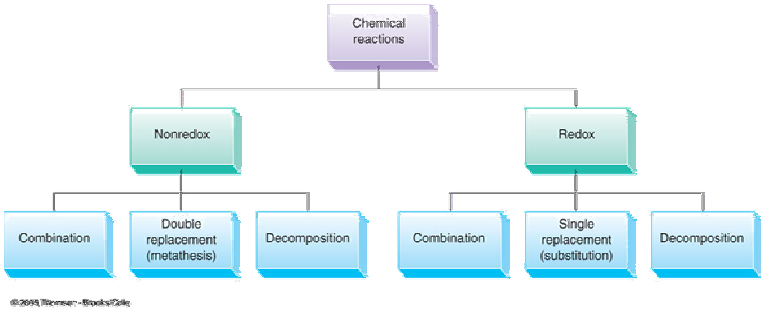
(s) for solid

(*l*) for liquid

(g) for gas

(aq) for a species in water

Characterization of Chemical Reactions:



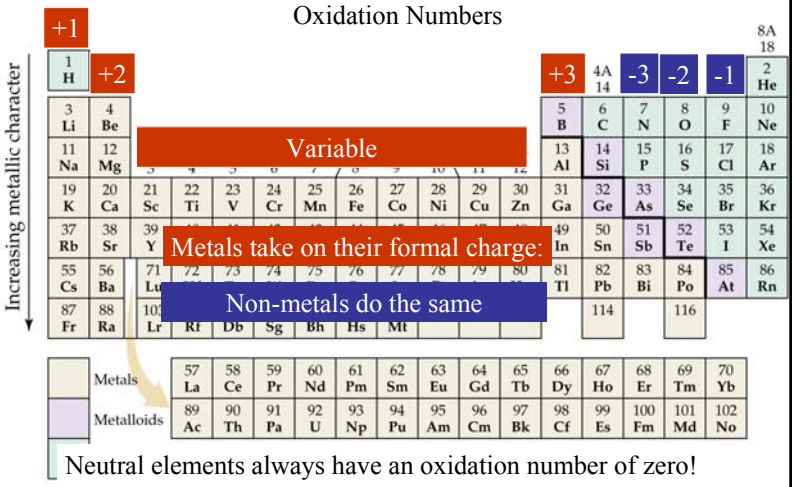
Reduction and Oxidation Reactions: (RedOx)

Reduction: elements or compounds that gain electrons during a reaction

Oxidation: elements or compounds that lose electrons during a reaction

Oxidation numbers: A bookkeeping tool used in chemistry to account for the exchange of electrons in a RedOx process.

Oxidation Numbers



Oxidation numbers

In the compound potassium bromate ($KBrO_3$), the oxidation number of bromine (Br) is?

The compound is neutral so the sum of the oxidation numbers should be zero.

+1 ? 3×(-2) = -6

$KBrO_3$

1+ ? + (-6) = 0

? = 5

Examples: Combination Reactions

Reaction of a metal with oxygen

calcium + oxygen

$Ca(s) + O_2(g) \longrightarrow CaO(s)$

How did I know to write CaO(s)

Ca forms a +2 cation, O forms a 2- anion

therefore: CaO(s) because all salts are solids under normal conditions.

$Ca(s) + O_2(g) \longrightarrow CaO(s)$

1 Ca & 2O's 1 Ca & 1O

Not balanced!

To balance, one must multiply each component of the reaction by coefficients that bring about balance.

One needs 2 O's on the right, multiply CaO by 2

$Ca(s) + O_2(g) \longrightarrow 2\,CaO(s)$

Now balance Ca with another "2"

$2\,Ca(s) + O_2(g) \longrightarrow 2\,CaO(s)$

Assigning Oxidation Numbers:

$$\overset{0}{2\,Ca(s)} + \overset{0}{O_2(g)} \longrightarrow \overset{+2\ \ -2}{2\,CaO(s)}$$

Ca goes form 0 to + 2, it has been ***oxidized***
O goes form 0 to – 2, it has been ***reduced***

Since O brings about the oxidation of Ca, it is the ***"Oxidizing agent"***

Since Ca brings about the reduction of O, it is the ***"Reducing agent"***

Metal & Non-metal: Transitions metals

For the case where a transition metal combines with a non-metal, one must be given the identity of the products.

Example: iron + oxygen

one can form: iron (II) oxide or iron (III) oxide

iron + oxygen forming iron (II) oxide

$$\overset{0}{2\ Fe(s)} + \overset{0}{O_2(g)} \longrightarrow \overset{+2\ \ -2}{2\ FeO(s)}$$

Reducing agent *oxidizing agent*

Recognizing a RedOx Reaction:

Whenever one reactant goes from a zero charge to a non zero charge, a RedOx reaction has occurred:

oxygen alone has a zero charge *oxygen in a compound has a –2 charge*

$$\overset{0}{CH_4(g)} + 2\,\overset{0}{O_2(g)} \rightarrow \overset{+4\ \ -2\ \ -2}{CO_2(g)} + 2\,H_2O(g)$$

Since the oxidation number of oxygen changed from 0 to –2 it was reduced!

If something was ***reduced***, something else was *oxidized*!

oxygen alone has a zero charge *oxygen in a compound has a –2 charge*

$$\overset{0}{CH_4(g)} + 2\,\overset{0}{O_2(g)} \rightarrow \overset{+4\ \ -2\ \ -2}{CO_2(g)} + 2\,H_2O(g)$$

CH_4

$? + 4\times(+1) = 0$

$? = -4$

CO_2

$? + 2\times(-2) = 0$

$? = +4$

Carbon goes from –4 to +4… it is oxidized!

Balancing ***REDOX*** reactions:

$$Fe + O_2 \rightarrow Fe_2O_3$$

oxidation #'s: 0 0 +3 -2

oxidation half reaction: $\{Fe \rightarrow Fe^{+3} + 3e^-\}$ x 4

reduction half reaction: $\{O_2 + 4e^- \rightarrow 2O^{2-}\}$ x 3

Balance electrons transferred then sum the half RXN's:

$$4Fe + 3O_2 + 12e^- \rightarrow 2Fe_2O_3 + 12e^-$$

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3$$