Chemistry 6A F2007

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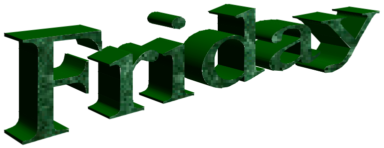

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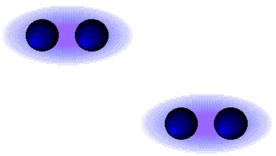
Dispersion Forces: Weakest of all intermolecular forces.

The forces that hold non–polar molecules together can be explained as follows:

Attractions arise from instantaneous, temporary dipoles formed due to electron motions.

The electron cloud of a molecule can be polarized to produce a short lived dipole that results in an attractive force.



Induced Dipole Forces:

Table 13.2 The Solubility of Some Gases in Water*

Gas	Molar Mass (g/mol)	Solubility at 20 °C (g gas/100 g water)†
H_2	2.01	0.000160
N_2	28.0	0.00190
O_2	32.0	0.00434

Degree to which electron cloud of an atom or molecule can be distorted is measured by its ***polarizability***.

As the electrons in a molecule become more loosely held and more spread out, the greater the degree of polarizbility a molecule has.

The explains the trend we see in solubility.

Dispersion Forces:

The Polarizability of a molecule is the ease with which an electron cloud can be distorted.

The larger the molecule (the greater the number of electrons) the greater polarizablity.

London dispersion forces increase as molecular weight increases.

London dispersion forces exist between all molecules.

London dispersion forces depend on the shape of the molecule.

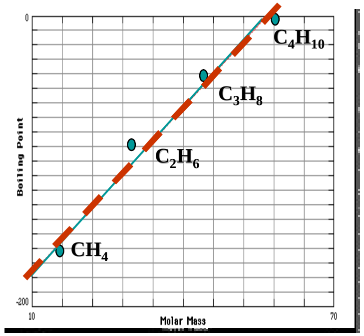
The greater the surface area available for contact, the greater the dispersion forces.

For molecules with the same relative shape, the forces scale with molar mass:

Higher molec. weight → larger induced dipoles.

Molecule	Boiling Point (°C)
CH_4 (methane)	- 161.5
C_2H_6 (ethane)	- 88.6
C_3H_8 (propane)	- 42.1
C_4H_{10} (butane)	- 0.5

Note linear relation between bp and molar mass.



Dispersion Forces:

As molar mass increases amongst non–polar molecules, bp and mp increase due to the increased polarizability of the molecule.

Halogen	mp (K)	bp (K)
F_2	53.5	85.0
Cl_2	172.2	238.6
Br_2	265.9	331.9
I_2	386.7	457.5

Chapter 5: Chemical Reactions:

A chemical equation describes the process of collections of compounds or atoms (Reactants) coming together in some manner to form new collections of compounds or atoms (Products).

Reactants ⟶ Products

What one starts with rearranges to form *new compounds*

Chemical Equations:

One more very important characteristic of chemical equations is that they must be ***balanced for mass***.

Total Mass of Reactants = Total Mass of Products

the number of atoms (A, B & C etc…) on the left = the number of atoms (A, B & C etc…) on the right

Balancing Chemical Equations

1: Write the unbalanced equation using correct chemical formulas for all reactants and products:

$$H_2 + O_2 \rightarrow H_2O$$

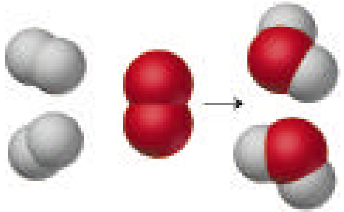
2:Find suitable coefficients to indicated how many formula units of each substance are required to balance the equation.

The 1 is implied so we omit it.

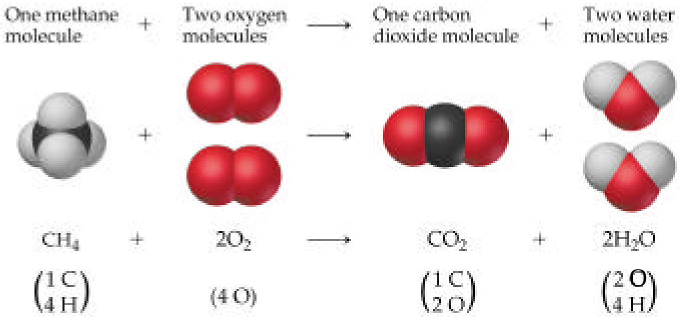
$$\underline{\ 2\ }H_2 + \underline{\ \ \ \ }O_2 \rightarrow \underline{\ 2\ }H_2O$$

The chemical equation for the formation of water can be visualized as two hydrogen molecules reacting with one oxygen molecule to form two water molecules:

$$2H_2 + O_2 \rightarrow 2H_2O$$



<u>Law of conservation of mass:</u> Matter cannot be lost in any chemical reactions.



<u>Balancing Combustion Reactions:</u> 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 (O₂)

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*

Stoichiometric balancing coefficients

They are the numbers in front of the chemical formulas.

They give the ***ratio*** of reactants and products.

ratio → ***Conversion factors***

The balancing coefficients allow us to convert between numbers of reactants and products.

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:

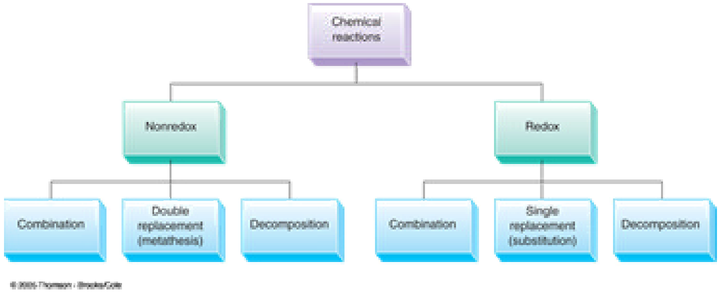
(s) for solid

(*l*) for liquid

(g) for gas

(aq) for a species in water

Characterization of Chemical Reactions:



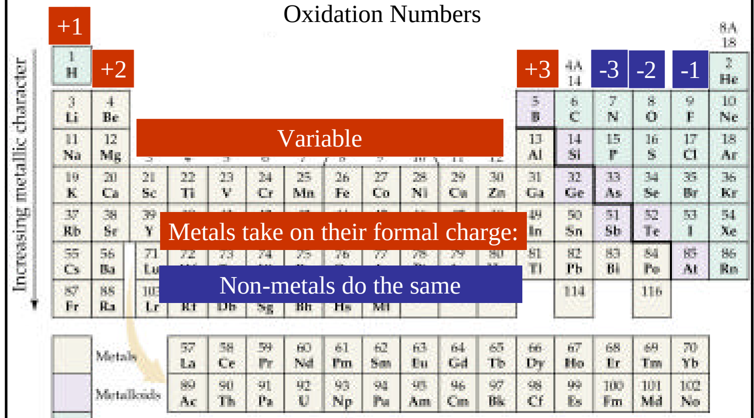
Reduction and Oxidation Reactions: (RedOx)

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



Neutral elements always have an oxidation number of zero!

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:

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

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

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

Reducing agent (Fe)   *oxidizing agent* (O_2)

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 → 0 (on O_2)

oxygen in a compound has a -2 charge → –2 (on CO_2), –2 (on H_2O)

$$CH_4(g) + 2\,\overset{0}{O_2}(g) \rightarrow C\overset{-2}{O_2}(g) + 2\,H_2\overset{-2}{O}(g)$$

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

$$\overset{-4}{CH_4}(g) + 2\,\overset{0}{O_2}(g) \rightarrow \overset{+4}{C}\overset{-2}{O_2}(g) + 2\,H_2\overset{-2}{O}(g)$$

CH_4

? + 4×(+1) = 0

? = –4

CO_2

? + 2×(–2) = 0

? = +4

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