Outline

- Oxidation Numbers
- Redox Reactions
- Balancing Redox Equations

Oxidation Numbers

An oxidation number is the charge that an atom “appears to have”… if electrons are assigned to more electronegative atom

HCl H is +1 and Cl is -1

Oxidation numbers represent number of electrons lost, gained, or unequally shared by an atom

An atom with oxidation number of…

+n has n fewer electrons in a molecule than as a free atom

-n has n more electrons in a molecule than as a free atom

The oxidation number…

…of atoms in elemental state is 0

…of monatomic ions is equal to charge of ions

…is +1 for Group I metals and +2 for Group II metals

…is usually +1 for hydrogen (except with metal hydrides)

…is usually -2 for oxygen (except with peroxides)

…is negative for the most electronegative atom in a compounds (and equal to charge of the ion)

The sum of oxidation numbers…

…in a compound is equal to zero

…in polyatomic ions is equal to charge of ion

Oxidation Number?

S_8	S = 0
S^{2-}	S = -2
Na_2S	S = -2
H_2S	S = -2
SF_6	S = +6
SO_3^{2-}	S = +4

More Oxidation Numbers!

CaH_2	Ca (+2); H (-1)
NaO	Na (+1); O (-1)
OF_2	O (+2); F (-1)
$KMnO_4$	K (+1); Mn (+7); O(-2)
$Al(OH)_4^-$	Al (+3); O (-2); H (+1)
$KClO_4$	K (+1); Cl (+7); O (-2)
H_2CO_3	H (+1); C (+4); O (-2)

Redox Reactions

A change in oxidation number implies loss or gain of electrons

Oxidation is the loss of electrons (increases oxidation number)

Reduction is the gain of electrons (decreases oxidation number)

$$\underset{0}{Cl_2} + \underset{-1}{2I^-} \rightarrow \underset{-1}{2Cl^-} + \underset{0}{I_2}$$

I^- undergoes oxidation… it is “oxidized”

Cl_2 undergoes reduction… it is “reduced”

Oxidation and reduction always accompany one another!

Oxidation sometimes accompanied with gain of oxygen or loss of hydrogen

Reduction sometimes accompanied with loss of oxygen or gain of hydrogen

$Pb \rightarrow PbO_2$	Pb oxidized
$CO \rightarrow CH_3OH$	CO reduced
$C_3H_4O \rightarrow C_3H_6O$	C_3H_4O reduced
$Fe \rightarrow Fe_2O_3$	Fe oxidized
$CH_4 \rightarrow CO_2$	CH_4 oxidized

Oxidation and reduction always accompany one another!

Substances that are oxidized are called the <u>reducing agents</u>… they cause something to be reduced

Usually: metals, H_2, elemental C

Substances that are reduced are called <u>oxidizing agents</u>… they cause something to be oxidized

Usually: halogens, O_2, $Cr_2O_7^{2-}$, MnO_4^-, HNO_3, H_2O_2

$Ni + F_2 \rightarrow NiF_2$	RA:	Ni	OA:	F_2
$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$	RA:	C	OA:	Fe_2O_3
$C_4H_8 + 6O_2 \rightarrow 4CO_2 + 4H_2O$	RA:	C_4H_8	OA:	O_2

Reactions in which atoms gain and lose electrons are called <u>oxidation-reduction</u> reactions (or a <u>Redox</u> reaction)

Oxidation numbers are used to keep track of electrons in redox reactions

In redox reactions, electrons gained by one element must equal electrons lost by another element!

e- lost in oxidation = e- gained in reduction

$$
\begin{array}{cccccccc}
Ni & + & F_2 & \rightarrow & Ni^{2+} & + & 2F^- & \rightarrow & NiF_2 \\
0 & & 0 & & +2 & & 2(-1) & &
\end{array}
$$

Balancing Redox Equations

Redox reactions balanced with the half-reaction method

1. Determine oxidation numbers

2. Write half-reactions for oxidation and reduction

3. Balance half-reactions with "MOHe"
   - M – miscellaneous atoms
   - O – oxygen atoms (with H_2O)
   - H – hydrogen atoms (with H^+)
   - e – electrons

4. Equalize electrons transferred

5. Combine half-reactions

Balance these redox reactions…

$Fe + Ag^{+} \rightarrow Fe^{3+} + Ag$

red: $3Ag^{+} + 3e^{-} \rightarrow 3Ag$

ox: $Fe \rightarrow Fe^{3+} + 3e^{-}$

redox: $\underline{Fe + 3Ag^{+} \rightarrow Fe^{3+} + 3Ag}$

$I^{-} + SO_4^{2-} \rightarrow H_2S + I_2$

red: $10H^{+} + SO_4^{2-} + 8e^{-} \rightarrow H_2S + 4H_2O$

ox: $8I^{-} \rightarrow 4I_2 + 8e^{-}$

redox: $\underline{8I^{-} + 10H^{+} + SO_4^{2-} \rightarrow H_2S + 4H_2O + 4I_2}$

Balance under basic conditions by adding OH^- to both sides…

$Cr_2O_3 + ClO^- \rightarrow CrO_4^{2-} + Cl^-$

red: $6H^+ + 6e^- + 3ClO^- \rightarrow 3Cl^- + 3H_2O$

ox: $5H_2O + Cr_2O_3 \rightarrow 2CrO_4^{2-} + 10H^+ + 6e^-$

redox: $2H_2O + Cr_2O_3 + 3ClO^- \rightarrow 2CrO_4^{2-} + 4H^+ + 3Cl^-$

with OH^-:

$4OH^- + 2H_2O + Cr_2O_3 + 3ClO^- \rightarrow 2CrO_4^{2-} + 4H_2O + 3Cl^-$

$\underline{4OH^- + Cr_2O_3 + 3ClO^- \rightarrow 2CrO_4^{2-} + 2H_2O + 3Cl^-}$