Slide 1

Oxidation-Reduction Reactions

RedOx Reactions

Slide 2

It’s barely chemistry…

In an oxidation-reduction reaction (abbreviated “redox” reaction), the only thing that changes is the “oxidation state” of the atoms involved.

Slide 3

What is an “oxidation state”?

The simplest way to think about an “oxidation state” is that it is the charge the atom has or could have if you separated it from the atoms it is bonded to.

Slide 4

Ionic compounds

For an ion or an ionic compound, the oxidation state is easy: it's the charge on the ion.

$FeCl_3$ – Fe must be +3 because there are 3 Cl^- ions stuck to it.

MnO_2 – Mn must be +4 because there are 2 O^{2-} stuck to it.

Slide 5

Covalent compounds

Here is where it is a little tricky. The atoms in a covalent compound don't have a real charge on them. BUT, they do have a potential charge if you pulled them all apart and the electrons they share get split up.

CO – carbon must be +2 since O wants to be -2
CO_2 – carbon must be +4 since there are 2 O that want to be -2 each

Slide 6

Rules for Oxidation states

1. Free element is 0 – including diatomics: Cu, Cl_2, O_2 are all ZERO.
2. An ion's charge is its oxidation state
3. The sum of the oxidation states of all atoms in a molecule equal the charge of the molecule.
4. Group 1a metals are always +1, Group 2a metals are always +2

Group VIIa nonmetals are usually -1
Group VIa nonmetals are usually -2

Slide 7

What's the oxidation state of the atoms?

$SrBr_2$

Br is a halogen – it must be -1 when bonded to a metal

Sr must be +2

SO_3

O is usually -2, which means S must be +6

Slide 8

What's the oxidation state of the atoms?

CO_3^{2-}

O is usually -2, which means C must be +4 since the entire molecule is -2

NO_3^-

O is usually -2, which means N must be +5 since the entire moecule is -1

Slide 9

Simplest Redox Reaction

$Fe^{2+} + Mn^{5+} \rightarrow Fe^{3+} + Mn^{2+}$

There is a transfer of electrons from the iron to the manganese resulting in the change in charge of each of them.

Slide 10

Balancing Redox Reactions

$Fe^{2+} + Mn^{5+} \rightarrow Fe^{3+} + Mn^{2+}$

Balancing redox reactions require the ELECTRONS to be balanced as well as the atoms.

$3\ Fe^{2+} + Mn^{5+} \rightarrow 3\ Fe^{3+} + Mn^{2+}$

(Chapter 18 discusses balancing in greater detail)

Slide 11

Recognizing redox reactions

All redox reactions involve 2 species that change oxidation state. It's a "donation" of electrons and, like acid/base chemistry, you can't have a donor without an acceptor!

$3\ Fe^{2+}$ +	$Mn^{5+} \rightarrow 3\ Fe^{3+} + Mn^{2+}$
Donor	Acceptor
Loses e-	Gains e-
Oxidized	Reduced
Reducing agent	Oxidizing Agent

Slide 12

Predicting Redox Reactions

We can't do it! (At least not until Chapter 18 ☺)

For now, we'll limit ourselves to recognizing them when we see them.

Slide 13

Combustion

Combustion – burning in the presence of oxygen – is often an example of a redox reaction:

$CH_4 + O_2 \rightarrow CO_2 + H_2O$

First thing you should notice:
Oxygen gets combined!
O_2 has a "0" oxidation state, it becomes "-2" when combined.
What else is changing oxidation state?
C – it is -4 in CH_4 and +4 in CO_2!

Slide 14

Recognizing Redox reactions

What gets oxidized, what gets reduced?

$4\ Li + O_2 \rightarrow 2\ Li_2O$

$Mg + Fe(NO_3)_2 \rightarrow Mg(NO_3)_2 + Fe$

Slide 15

Clicker Question

Which of the following is a redox reaction:

A. $Li_{(s)} + O_{2\ (g)} \rightarrow Li_2O_{\ (s)}$

B. $Pb(NO_3)_{2\ (s)} + Na_2SO_{4\ (s)} \rightarrow PbSO_{4\ (s)} + 2\ NaNO_{3\ (s)}$
C. $Mg_{(s)} + Br_{2\ (l)} \rightarrow MgBr_{2\ (s)}$
D. A and C
E. A, B, and C.