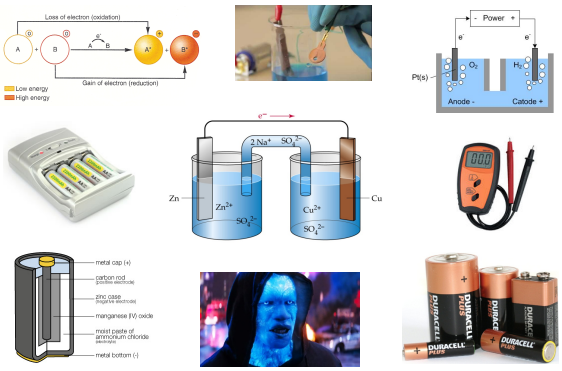
Intro to Redox



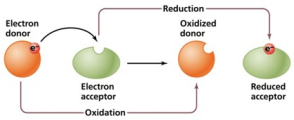
Electrochemistry

Study of the interchange of chemical and electrical energy

- Electron transfer reactions are called oxidation-reduction reactions or REDOX reactions
- Electrochemical processes that result in the generation of an electric current (electricity) or can be caused by imposing an electric current

Redox Reactions

- **OXIDATION NUMBER:** Assigned charge on an atom
- **OXIDATION:** Loss of electrons (increase in oxidation number or charge… "+")
- **REDUCTION:** Gain of electrons (decrease in oxidation number or charge… "-")



Redox Reactions

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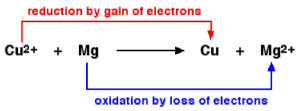
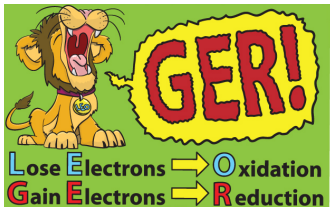


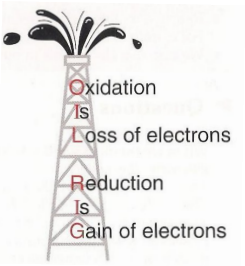
Redox Reactions

LEO THE LION GOES GER!



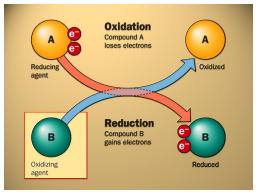
Redox Reactions

OIL RIG!



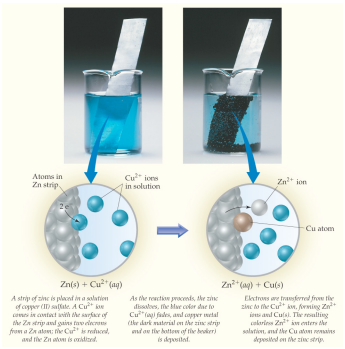
Redox Reactions

- **OXIDIZING AGENT:** electron acceptor… species that is reduced (an agents facilitates something / ex: travel agent)
- **REDUCING AGENT:** electron donor… species that is oxidized



Redox Reactions

Placing a piece of Zn in a solution with Cu^{2+} ions… Make note of what is taking place at the molecular level!



Redox Reactions

- Oxidation and reduction go hand in hand… need one to have the other
- Cannot have 2 oxidations or 2 reductions in the same equation
- Written as two **HALF-REACTIONS** (one for oxidation and one for reduction)

$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$ Overall

$2\ Mg \rightarrow 2\ Mg^{2+} + 4\ e^-$ OXIDATION

$O_2 + 4\ e^- \rightarrow 2\ O^{2-}$ REDUCTION

Oxidation Numbers

Assigned charge on an atom

1) Elements not bonded to another different element have an oxidation number of **ZERO** (ex: Na, Fe, O_2, N_2)

2) In monatomic ions, oxidation number is equal to the charge on the ion (ex: $Li^+ = +1$, $Fe^{3+} = +3$, $P^{3-} = -3$)

3) Oxidation number of oxygen is **USUALLY** -2… in H_2O_2 and O_2^{2-} it is -1

Oxidation Numbers

Assigned charge on an atom

4) Oxidation number of hydrogen is +1 **EXCEPT** when bonded to metals in binary compounds then it is -1 (ex: CaH_2, LiH)

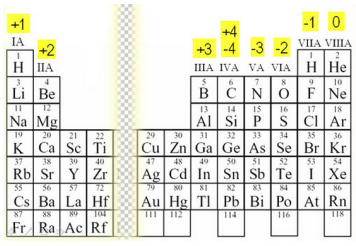
5) Group IA metals are always +1, IIA always +2, etc… remember transition metals vary

6) Oxidation numbers of a molecule must add up to **ZERO** or add up to the charge on a polyatomic ion

Ex: HCO_3^-

Oxidation Numbers

Assigned charge on an atom



Oxidation Numbers

Assigned charge on an atom

Examples

CO_2 CO C CH_4 H_2 Cl_2 HCl
(+4)(-2) (+2)(-2) (0) (-4)(+1) (0) (0) (+1)(-1)

$NaCl$ $KMnO_4$ ZnO CaF_2
(+1)(-1) (+1)(+7)(-2) (+2)(-2) (+2)(-1)

O_2 O_3 HO_2 H_2O_2 H_2O OF_2
(0) (0) (+1)(-1/2) (+1)(-1) (+1)(-2) (+2)(-1)

Practice

- **EXAMPLE:**

Identify which is being oxidized and which is being reduced in the following.

$2\ Ca(s) + O_2(g) \rightarrow 2\ CaO(s)$

$Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

$2\ Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$

Writing a Redox Reaction

1) Write the NET IONIC equation

2) Assign oxidation numbers

3) Determine what is being oxidized and reduced

- **EXAMPLE:** Lead foil is immersed in silver nitrate

$Pb(s) + 2\ Ag^+(aq) \rightarrow Pb^{2+}(aq) + 2\ Ag(s)$
0 **+1** **+2** **0**

Balancing a Redox Reaction

1) Write the half reactions

2) Balance the non H's and O's first

3) Use H_2O to balance the O's and use H^+ to balance the H's

4) Use e- to balance the charges

5) Get common multiple for e- and multiply

6) Cancel and add up the half reactions (e- must cancel as well as all like species)… Verify # of atoms and charges are balanced

7) In basic solutions, add OH^- to BOTH sides for every H^+ in final equation… make H_2O

Balancing a Redox Reaction

- **EXAMPLE:**

Write the balanced equation for the oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution.

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

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$

STEP #1

Balancing a Redox Reaction

- **EXAMPLE:**

Write the balanced equation for the oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution.

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

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$

STEP #2

Balancing a Redox Reaction

- **EXAMPLE:**

Write the balanced equation for the oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution.

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

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

STEP #3

Balancing a Redox Reaction

- **EXAMPLE:**

Write the balanced equation for the oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution.

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

$14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

STEP #3

Balancing a Redox Reaction

- **EXAMPLE:**

Write the balanced equation for the oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution.

$Fe^{2+} \rightarrow Fe^{3+} + 1\ e^-$

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

STEP #4

Balancing a Redox Reaction

- **EXAMPLE:**

Write the balanced equation for the oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution.

$6\ Fe^{2+} \rightarrow 6\ Fe^{3+} + 6\ e^-$

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

STEP #5

Balancing a Redox Reaction

- **EXAMPLE:**

Write the balanced equation for the oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ to Cr^{3+} in acidic solution.

$6\ Fe^{2+} \rightarrow 6\ Fe^{3+} + \cancel{6\ e^-}$

$\cancel{6\ e^-} + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

$14\ H^+ + Cr_2O_7^{2-} + 6\ Fe^{2+} \rightarrow 6\ Fe^{3+} + 2\ Cr^{3+} + 7\ H_2O$

STEP #6 **Are all atoms and charges balanced?**

Practice

- **EXAMPLES:**

Write the balanced equation for the following reaction in **BASIC** solution:

$ClO_3^- + SO_2 \rightarrow Cl_2 + SO_4^{2-}$

Write the balanced equation for the following reaction in **ACIDIC** solution:

$MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$

Practice

- **EXAMPLE:**

  Write the balanced equation for the following reaction in **BASIC** solution:

  $CN^- + MnO_4^- \rightarrow CNO^- + MnO_2$