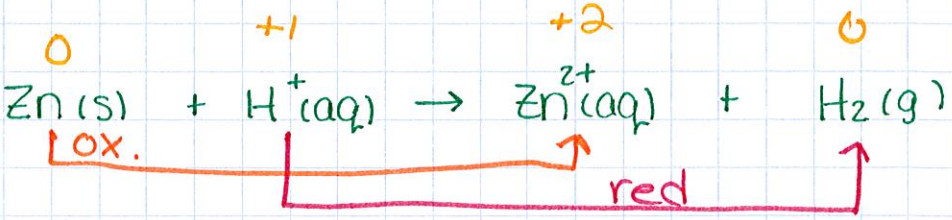
Ch 20 - Electrochemistry

1. Electrochemical Reactions - e^- are transferred from one species to another (redox reactions)

A. How do you know a redox reaction has occurred?

1st assign oxidations
2nd did one species get oxidized (lose e^-) and another get reduced (gain e^-)?

(Ex) :



which species got oxidized? Zn
which species got reduced? H

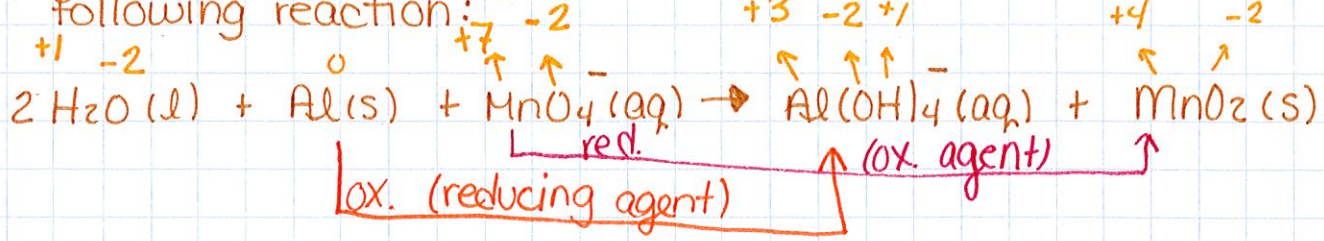
(H) 1. oxidizing agent - substance that gets reduced/cause oxidation in another species

(Zn) 2. reducing agent - substance that gets oxidized/cause reduction in another species

(Ex)

not redox! b/c water isn't ionic, it's covalent so the e^- are shared not transferred.

Ex (1) Identify the oxidizing and reducing agents in the following reaction:



Ex (2) $\overset{6+\ -2-}{Cr_2O_7(aq)} + \overset{-2+1-}{2OH(aq)} \rightarrow \overset{6+\ -2\ 2-}{2CrO_4(aq)} + \overset{+1\ -2}{H_2O(l)}$

$12 - 14 = -2$; $6 - 8$

not redox

(3) $\overset{0}{O_3(g)} + \overset{+2\ -2}{NO(g)} \rightarrow \overset{0}{O_2(g)} + \overset{+4\ -2}{NO_2(g)}$

oxidized (red. agent)

reduced (ox. agent)

B. Refresher on Assigning Oxidation Numbers

	Ox. #
(1) Elements in their elemental form	$\emptyset$
(2) monatomic ions	same as charge
(3) Nonmetals tend to have negative oxidation numbers	
(a) Oxygen	-2
(b) Peroxide	-1
(c) Hydrogen when bonded to a metal	-1
(d) Hydrogen when bonded to a nonmetal	$+1$
(e) Fluorine	-1
(f) other Halogens except they can be + in polyatomic ions	-1

(4) Neutral compounds - sum of oxidation numbers is $\emptyset$.
(5) Polyatomic ions - sum of oxidation numbers is charge.

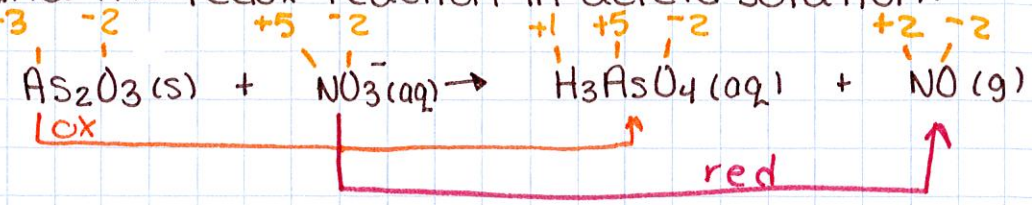
C. Balancing Redox Reactions (half-reaction method)
(Acidic Solution)

- involves treating the oxidation and reduction as 2 separate processes, balancing the half reactions, and then combining them to attain the balanced equation for the overall reation.

Steps

1. Assign oxidation numbers to determine what is oxidized and what is reduced.
2. write the oxidation and reduction half reactions
3. Balance each half reaction
   a. Balance elements other than H and O.
   b. Balance O by adding H_2O.
   c. Balance H by adding H^+.
   d. Balance charge by adding e^-.
4. multiply half reactions by integers so that e^- gained is same as e^- lost
5. Add half reactions, subtracting things that appear on both sides.
6. make sure the equation is balanced by mass.
7. make sure the equation is balanced by charge.

Balance the redox reaction in acidic solution.



$5H_2O + 3As_2O_3(s) \rightarrow 6H_3AsO_4(aq) + 4H^+ + 4e^-$

$3e^- + 4H^+ + 4NO_3^-(aq) \rightarrow 4NO(g) + 2H_2O$

$7H_2O + 3As_2O_3 + 4H^+ + 4NO_3^- \rightarrow 6H_3AsO_4 + 4NO$

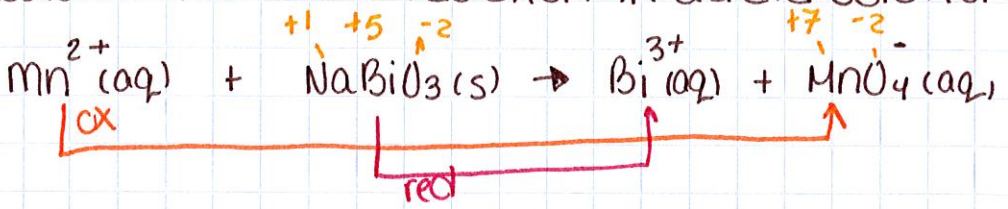
Ex (4) Balance this redox reaction in acidic solution.

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

$2Cl^- \rightarrow 3Cl_2 + 2e^-$

$14H^+ + Cr_2O_7^{2-} + 6Cl^- \rightarrow 2Cr^{3+} + 7H_2O + 3Cl_2$

Ex (5) Balance this redox reaction in acidic solution.



$$4H_2O + 3Mn^{+2} \rightarrow 3MnO_4^- + 8H^+ + 5e^-$$

$$+ 5e^- + 6H^+ + 5NaBiO_3 \rightarrow 5Bi^{3+} + 5Na + 3H_2O$$

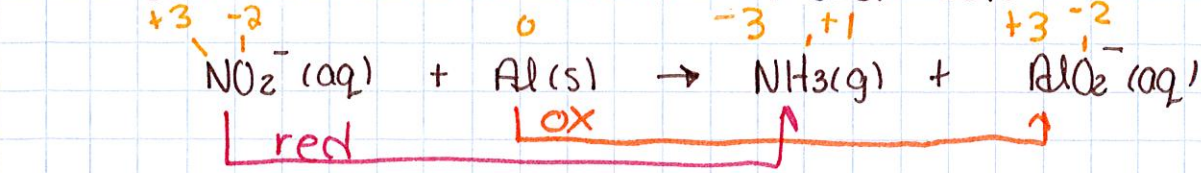
$$3Mn^{2+} + 4H_2O + 5NaBiO_3 \rightarrow 3MnO_4^- + 5Bi^{3+} + 5Na + 3H_2O$$

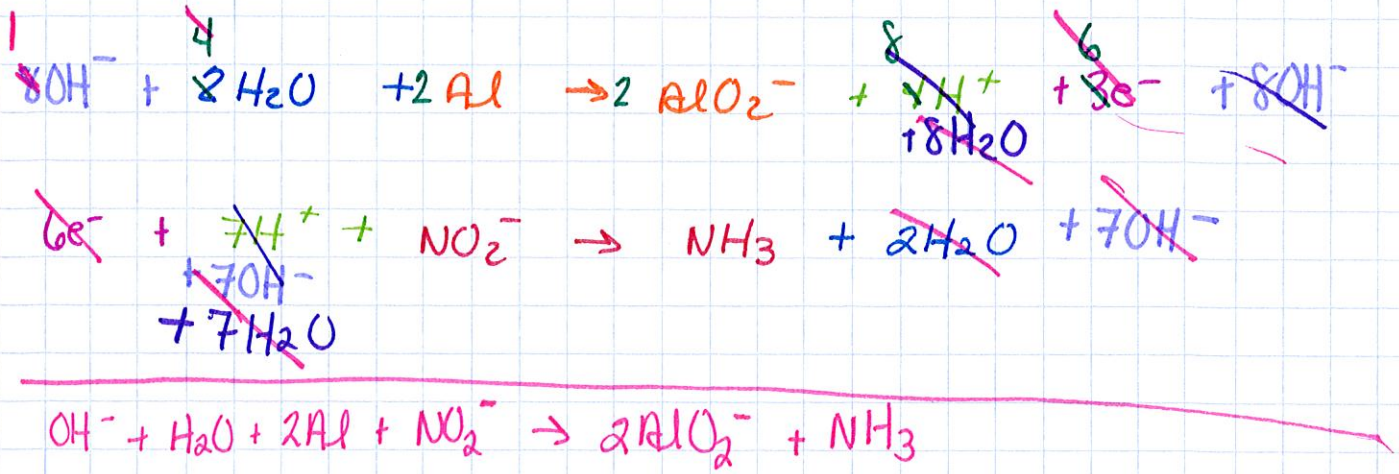
D. Balancing Redox Reactions in Basic Solution

Steps

1. Balance the reaction as if it were in acidic solution.
2. Add OH^- to each side to "neutralize" the H^+ in the equation and create H_2O in its place.
3. You might have to subtract H_2O from each side.

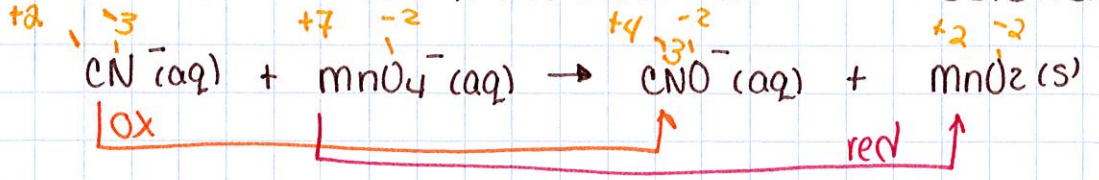
Balance the redox reaction in basic solutions.

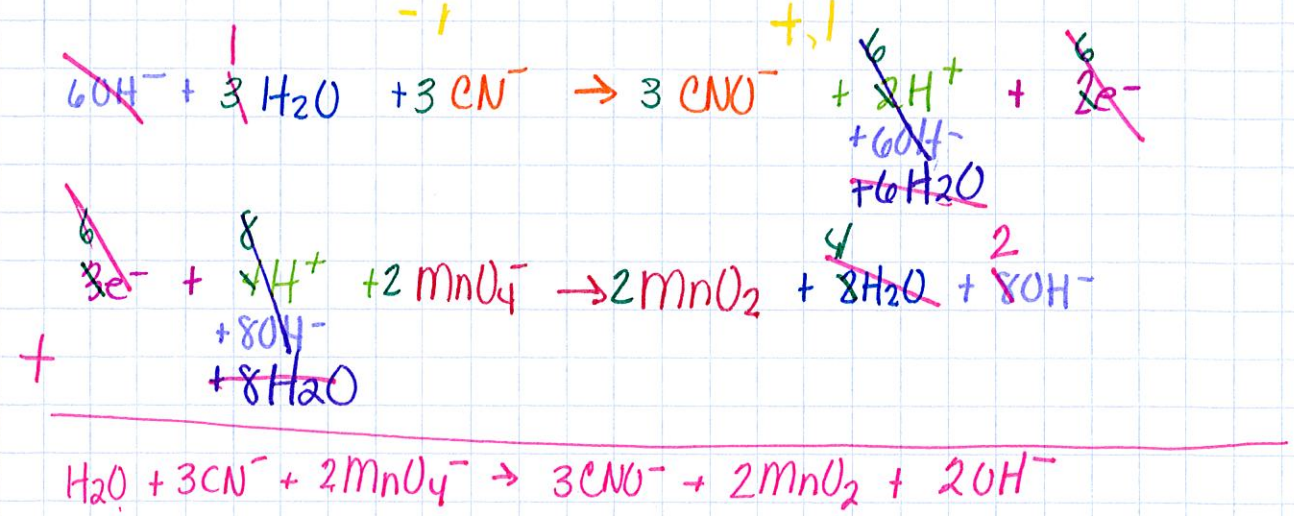




$OH^- + H_2O + 2Al + NO_2^- \rightarrow 2AlO_2^- + NH_3$

Ex (6) Balance this redox reaction in basic solution.





$H_2O + 3CN^- + 2MnO_4^- \rightarrow 3CNO^- + 2MnO_2 + 2OH^-$

HW p 884-885 # 16, 21, 22

Ex (7) Balance this redox reaction in basic solution

$\overset{+7}{Mn}\overset{-2}{O_4^-}(aq) + \overset{-2}{S^{2-}}(aq) \rightarrow \overset{+2}{Mn}\overset{-2}{S}(s) + \overset{0}{S}(s)$

red: MnO_4^- → MnS (+7 → +2)

ox: S^{2-} → S (−2 → 0)

$7\,S^{2-} \rightarrow 7S + 14e^-$

$14e^- + 16H_2O + 8H^+ + 2MnO_4^- \rightarrow 2MnS + 8H_2O + 16OH^-$

$8H^+ + 16OH^- \rightarrow 8H_2O$

$7S^{2-} + 8H_2O + 2MnO_4^- \rightarrow 7S + 2MnS + 16OH^-$

Ex (8) Balance this redox reaction in basic solution.

$\overset{0}{Cl_2}(g) \rightarrow \overset{-1}{Cl^-}(aq) + \overset{+1}{OCl^-}(aq)$

red: Cl_2 → Cl^- (0 → −1)

ox: Cl_2 → OCl^- (0 → +1)

$4OH^- + 2H_2O + Cl_2 \rightarrow 2OCl^- + 4H^+ + 2e^-$

$4OH^- \rightarrow 4H^+ + 4H_2O$

$2e^- + Cl_2 \rightarrow 2Cl^-$

$4OH^- + 2Cl_2 \rightarrow 2OCl^- + 2H_2O + 2Cl^-$