

ELECTROCHEMISTRY – 1

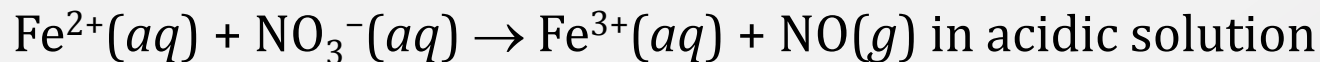
REDOX EQUATIONS

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VOLTAIC CELLS

- A **voltaic cell** employs a spontaneous oxidation–reduction reaction as a source of energy.
- The redox reactions studied here are electron transfers: oxidation – loss of electrons and reduction – gain of electrons.
- It separates the reaction into two half–reactions, physically separating one half–reaction from the other.
- Our first step in studying electrochemical cell is to balance its oxidation–reduction reaction in an acidic or basic solutions.
- The electron transfer is studied so the hydronium ion, $\text{H}_3\text{O}^+(\text{aq})$, will be represented by its simpler notation, $\text{H}^+(\text{aq})$. Only the notation, not the chemistry, is different.

A REDOX EQUATION



We determine the oxidation number of Fe and N in each substance.

Oxidation: $\text{Fe}^{2+}(\text{aq})$ is oxidized from +2 to +3 in $\text{Fe}^{3+}(\text{aq})$.

Reduction: N in $\text{NO}_3^{-}(\text{aq})$ is reduced from +5 to +2 in $\text{NO}(\text{g})$.

The reaction is in acidic conditions (HNO_3).

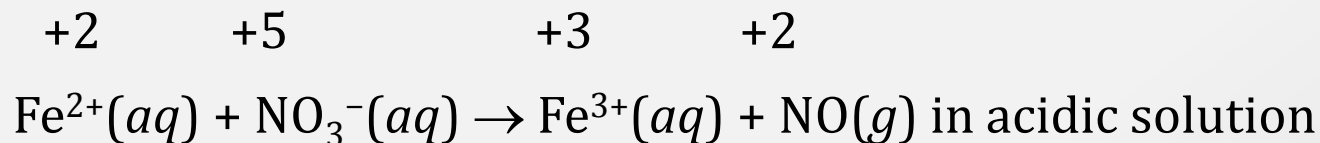
Then we proceed to balance the equation....next few slides.

BALANCING REDOX REACTIONS – IN ACID

To balance the reaction there is a five step process:

- 1) Balance all atoms other than H and O.
- 2) Balance O using H_2O on the opposite side.
- 3) Balance H using H^+ on the opposite side.
- 4) Balance the charges on both sides using electrons on opposite sides.
(remember for oxidation electrons will be on the product side and for reduction they will be on the reactant side)
- 5) Multiply the entire oxidation or reduction equation to have the same number of electrons for both equations.
- 6) Add the two redox reactions – that should nullify the electrons.

EXAMPLE: BALANCING REDOX EQUATION



Oxidation half-reaction



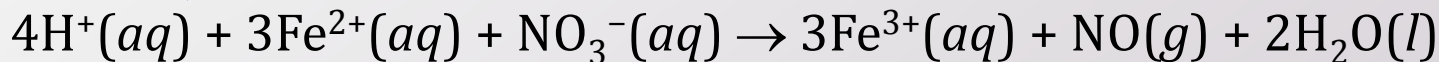
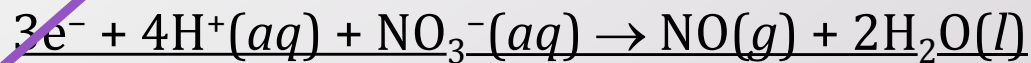
Reduction half-reaction

First we balance N, then O: $\text{NO}_3^{-}(\text{aq}) \rightarrow \text{NO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Next we balance H: $4\text{H}^{+}(\text{aq}) + \text{NO}_3^{-}(\text{aq}) \rightarrow \text{NO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Then we balance e^{-} : $3\text{e}^{-} + 4\text{H}^{+}(\text{aq}) + \text{NO}_3^{-}(\text{aq}) \rightarrow \text{NO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Combine the two equations



BALANCING REDOX REACTIONS – IN BASE

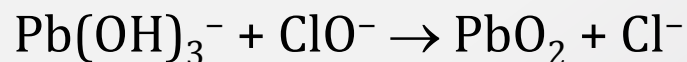
- All the rules are the same as in the acidic conditions.

Then

1. Add one OH^- to both sides for each H^+ .
2. When H^+ and OH^- occur on the same side, combine them to form H_2O .
3. Cancel water molecules that occur on both sides.

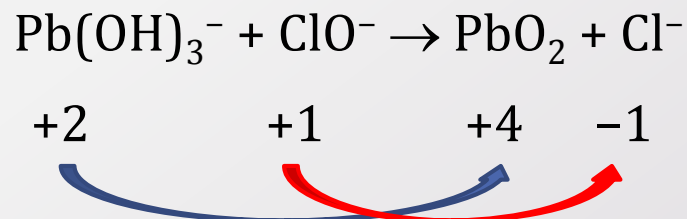
EXAMPLE: BALANCING REDOX EQ IN BASE

Lead(II) ion, Pb^{2+} , yields the plumbite ion, $\text{Pb}(\text{OH})_3^-$, in basic solution. In turn, this ion is oxidized in basic hypochlorite solution, ClO^- , to lead(IV) oxide, PbO_2 . Balance the equation for this reaction using the half-reaction method. The skeleton equation is



Solution:

First determine redox..



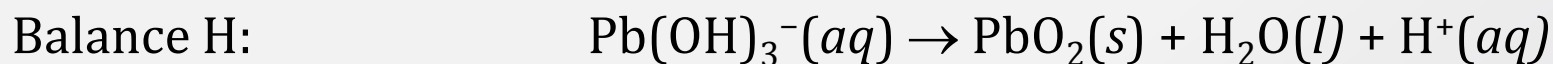
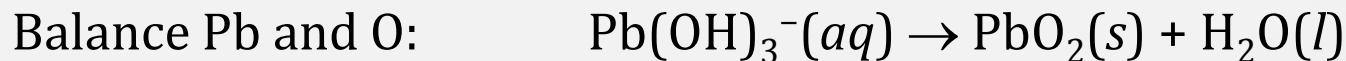
Pb is oxidized from +2 to +4.

Cl is reduced from +1 to -1.

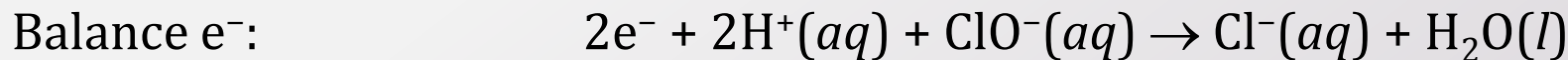
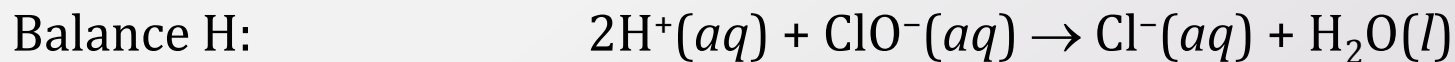
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EXAMPLE: REDOX EQ IN BASE..CONTD

Oxidation half-reaction



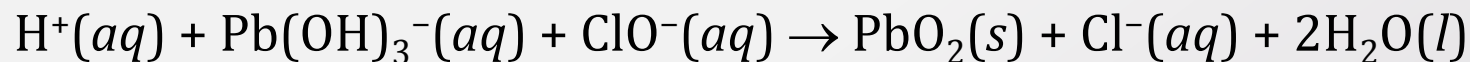
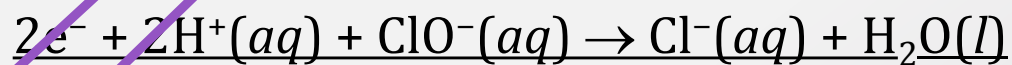
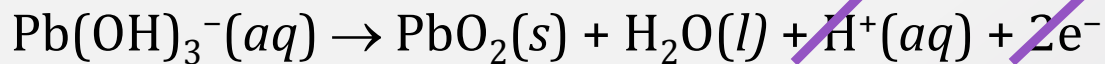
Reduction half-reaction



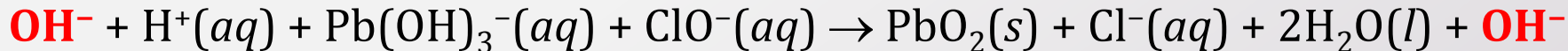
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EXAMPLE: REDOX EQ IN BASE..CONTD

Combine the half-reactions.



To convert to basic solution, we add OH^- to each side, converting H^+ to H_2O .



Finally, we cancel one H_2O that is on both sides.



KEY CONCEPTS

- Balancing redox equations in acid and base medium