## Overview of Chapter 20

Electrochemistry

- Oxidation/Reduction (Redox) Reactions
- Electrochemical Cells
- E° and E
- E° and ΔG
- E° and K

## Topics to consider:

- Balance redox reactions

## Terminology for Redox Reactions

- OXIDATION—loss of electron(s) by a species; increase in oxidation number.
- REDUCTION—gain of electron(s); decrease in oxidation number.
- OXIDIZING AGENT—electron acceptor; species is reduced.
- REDUCING AGENT—electron donor; species is oxidized.

## OXIDATION-REDUCTION REACTIONS

Direct Redox Reaction

Oxidizing and reducing agents in direct contact.

$Cu(s) + 2\ Ag^{+}(aq) \text{--->} Cu^{2+}(aq) + 2\ Ag(s)$

## OXIDATION-REDUCTION REACTIONS

Indirect Redox Reaction

A battery transfers electrons through an external wire from the reducing agent to the oxidizing agent.

## Oxidation States of Atoms in Elements and Compounds

**IN ELEMENTS**

| | | Oxidation Number: | Examples: |
|---|---|---|---|
| 1. | **Atoms in their elemental state** | =0 | Fe, $H_2$, $O_2$ |
| 2. | **Monatomic ions** | =charge | $F^{1-}$, $Na^{1+}$, $Fe^{3+}$ |

**IN COMPOUNDS**

| | | | |
|---|---|---|---|
| 3. | **Group 1A** | =+1 | NaCl, $KNO_3$ |
| 4. | **Group 2A** | =+2 | MgO |
| 5. | **Fluorine** | =-1 | HF, ClF |
| 6. | **Hydrogen** | =+1 | $H_2O$ |
| 7. | **Oxygen** | =-2 | $SO_2$, $HClO_4$ |
| 8. | **Group 7A** | =-1 | HCl |
| 9. | **Group 6A** | =-2 | $PbS_2$ |

The sum of all oxidation numbers of all elements = charge on substance

## Balancing Oxidation-Reduction Reactions

1. Assign oxidation numbers.

2. Separate into oxidation and reduction half reactions.

3. Balance each half reaction using the following steps:
   a. Balance all elements except oxygen or hydrogen.
   b. Balance oxygen by adding $H_2O$.
   c. Balance hydrogen by adding H+.
   d. Balance charge by adding electrons:
      Electrons go on the RIGHT (product side) for OXIDATION reactions.
      Electrons go on the LEFT (reactant side) for REDUCTION reactions.
   e. In BASIC solution, do this additional step:
      For every H+, add OH- to BOTH sides of the reaction.
      Combine H+ + OH- into $H_2O$.
      Cancel out any waters that appear on both sides.

   You should now have a balanced half reaction.

4. Multiply balanced half reactions so an equal number of electrons are consumed and produced.

5. Add together half reactions.

6. Clean up. Combine identical substances and reduce coefficients to the lowest terms.

7. CHECK! Atom and charge must balance.

## Balancing Equations

$$Cu + Ag^+ \rightarrow Cu^{2+} + Ag$$

## Balancing Equations

Step 1: Assign oxidation numbers.

Cu and Ag = 0
$Cu^{2+}$ = +2
$Ag^+$ = +1

Step 2: Divide the reaction into half-reactions, one for oxidation and the other for reduction.

Ox $Cu \text{ --->} Cu^{2+}$
Red $Ag^+ \text{ --->} Ag$

Step 3: Balance each for mass and charge.
(mass already done in this case)

Ox $Cu \text{ --->} Cu^{2+} + 2e^-$
Red $Ag^+ + e^- \text{ --->} Ag$

## Balancing Equations

Step 4: Multiply each half-reaction by a factor so that the number of electrons on each side is the same.

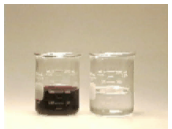
Oxidation $Cu \text{ --->} Cu^{2+} + 2e^-$
Reduction $2\,Ag^+ + 2\,e^- \text{ --->} 2\,Ag$

Step 5: Add half-reactions to give the overall equation.

$$Cu + 2\,Ag^+ \text{ --->} Cu^{2+} + 2Ag$$

The equation is now balanced for both charge and mass.

## Balancing Equations for Redox Reactions

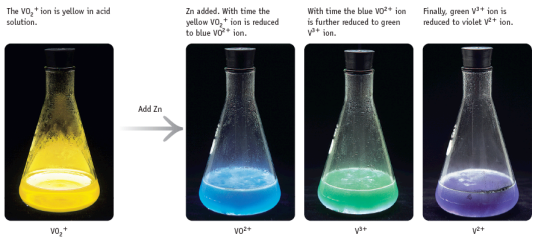


$$MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$$

**Mn = +7 Fe = +2 Mn = +2 Fe = +3**

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

## Reduction of $VO_2^+$ with Zn



## Balancing Equations

Balance the following in acid solution—

$VO_2^+ + Zn \text{ ---> } VO^{2+} + Zn^{2+}$

Step 1: Identify the oxidation states

V in $VO_2^+$ = +5 V in $VO^{2+}$ = +4

Zn = 0 $Zn^{2+}$ = +2

Step 2: Write the half-reactions

Ox $Zn \text{ ---> } Zn^{2+}$

Red $VO_2^+ \text{ ---> } VO^{2+}$

Step 3a: Balance each half-reaction for mass. Add $H_2O$ on O-deficient side and add $H^+$ on other side for H-balance.

Ox $Zn \text{ ---> } Zn^{2+}$

Red **$2\ H^+ + VO_2^+ \text{ ---> } VO^{2+} + H_2O$**

## Balancing Equations

Step 3b: Add electrons to balance half-reactions for charge.

Ox $Zn \text{ ---> } Zn^{2+} + 2e^-$

Red $e^- + 2\ H^+ + VO_2^+ \text{ ---> } VO^{2+} + H_2O$

Step 4: Multiply by an appropriate factor.

Ox $Zn \text{ ---> } Zn^{2+} + 2e^-$

Red $2e^- + 4\ H^+ + 2\ VO_2^+ \text{ ---> } 2\ VO^{2+} + 2\ H_2O$

Step 5: Add *balanced* half-reactions

$Zn + 4\ H^+ + 2\ VO_2^+ \text{ ---> } Zn^{2+} + 2\ VO^{2+} + 2\ H_2O$