

Chapter 20 - Electron Transfer Reactions (Electrochemistry)

Recall from Chapter 5:

oxidation: an increase in oxidation number or a process in which e^- are lost

reduction: a decrease in oxidation number or a process in which e^- are gained

oxidizing agent: a substance that oxidizes another substance; it itself gets REDUCED

reducing agent: a substance that reduces another substance; it itself gets OXIDIZED

In a redox reaction, both oxidation and reduction occur at the same time.

Oxidation Number Rules:

(1) The oxidation number of any free element is $\emptyset$.

(2) For a compound, the sum of all the oxidation numbers = $\emptyset$.

(3) For an ion, the sum of the oxidation numbers is the charge on the ion.

(4) For specific elements

IA metals +1

IIIA +3

IIA +2

oxygen -2 (except for H_2O_2 and a few others)

(5) For simple binary IONIC compounds: VA -3; VIA -2 VIIA -1
(metal + nonmetal)

Examples:

$NaCl$ Na +1 Cl -1

$NaClO$ Na +1 O -2 so Cl must be +1, $+1 + x + -2 = 0$

Fe^{3+} Fe +3

$Cr_2O_7^{2-}$ set up algebra $2x + 7(-2) = -2$

$$2x = +12$$

$$x = +6$$

NH_4^+ $x + 4(+1) = +1$; $x = -3$

Note on net ionic equations - Chapter 5

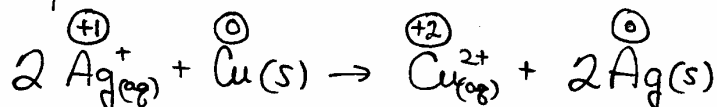
Formula Unit Equation $2\text{AgNO}_3(\text{aq}) + \text{Cu}(\text{s}) \rightarrow \text{Cu}(\text{NO}_3)_2(\text{aq}) + 2\text{Ag}(\text{s})$
 then break up strong electrolytes (strong acids, strong bases, soluble salts)

Total Ionic Equation $2\text{Ag}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{Ag}(\text{s})$
 then cancel out spectator ions (ions identical on both sides of equation)

Net Ionic Equation $2\text{Ag}^+(\text{aq}) + \text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$

Note: mass and charges balance.

This is a redox reaction: It can be recognized as such after assigning oxidation numbers to all elements.

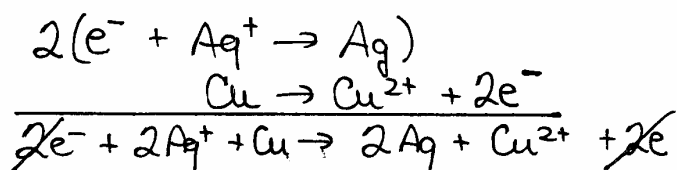


reduction: $\overset{+1}{\text{Ag}}^+ \rightarrow \overset{0}{\text{Ag}}$ $\therefore \text{Ag}^+$ is the oxidizing agent
 oxidation: $\overset{0}{\text{Cu}} \rightarrow \overset{+2}{\text{Cu}}^{2+}$ Cu is the reducing agent

These two reactions are called "half-reactions." To balance them charge-wise, we need to add electrons to one side or the other.

reduction half-rxn: $\text{e}^- + \text{Ag}^+ \rightarrow \text{Ag}$
 oxidation half-rxn: $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$

But the electrons involved in oxidation must equal the electrons involved in reduction, so we multiply the reduction half rxn by 2.



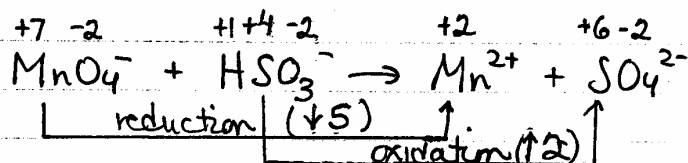
And the equation is balanced!

I'm going to show you how to balance redox net ionic equations using a variation of the half-rxn method.

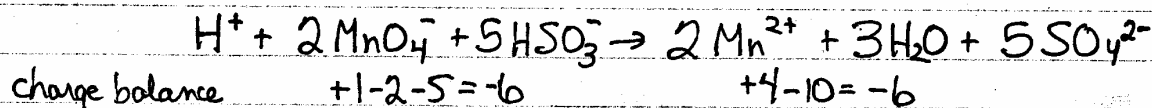
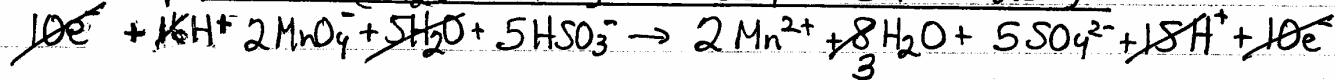
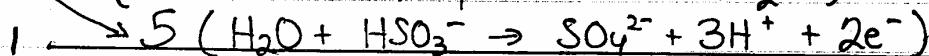
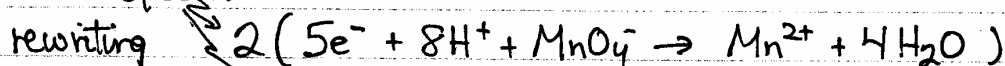
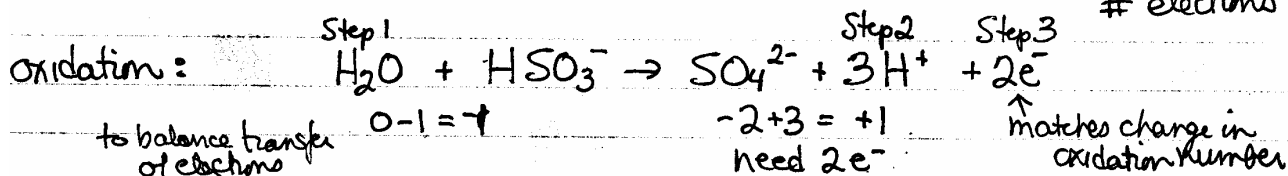
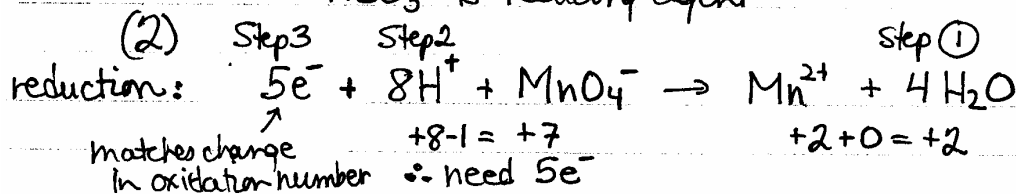
Note: in acidic solution, you can add as many H^+ or H_2O as you need
in basic solution, you can add as many OH^- or H_2O as you need.

Balance in acidic solution: $MnO_4^- + HSO_3^- \rightarrow Mn^{2+} + SO_4^{2-}$

(1) Assign oxidation numbers:



$\therefore MnO_4^-$ is oxidizing agent
 HSO_3^- is reducing agent



Note: In acidic solution:

- Step 0 - balance non-O/H elements
- Step 1 - balance O with H_2O
- Step 2 - balance H with H^+
- Step 3 - balance charge with e^-

