

In addition to understanding oxidation numbers to the point where they are second nature, you also have to be very good at balancing redox reactions. When you are given a reaction, you balance it, assign every atom an oxidation number and determine who is oxidized (loses electrons which are negatively charged so the oxidation number increases) and who is reduced (gains electrons so the oxidation number increases).

Here's a pretty good (quick and painless) Redox balancing technique

Step 1. Identify who is oxidized and who is reduced and break the reaction into two half-reactions.

- One for the oxidation $\longrightarrow$ The oxidation half-reaction
- One for the reduction $\longrightarrow$ The reduction half-reaction

Step 2. Do all the rest of these steps

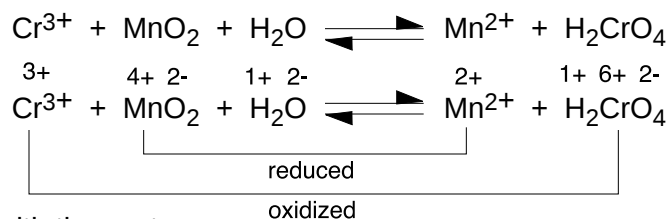
In acid solution

1. Balance all atoms other than H and O
2. Balance O with H_2O
3. Balance H with H^+
4. Balance charge (both sides have to have the same net charge) by adding electrons, e^-
5. Add the two half-reactions so that the e^- 's cancel algebraically
6. Cancel other species that appear on both sides of the equation
7. Check and make sure everybody's balanced

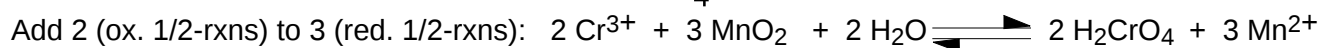
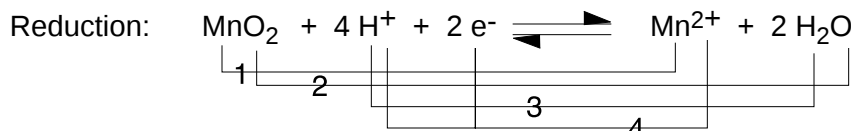
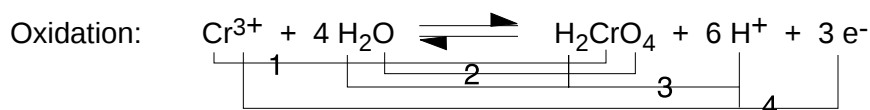
In basic solution

1. Do everything exactly the same as in acid solution (steps 1 to 5)
2. For each H^+ that's left, add an OH^- to both sides of the equation
3. On the side with H^+ , add the H^+ and the OH^- together to get H_2O
4. These should algebraically cancel with the H_2O that was already on the other side
5. Check and make sure everything's balanced

For example: Balance this reaction



2. It's in acid solution (H_2CrO_4) so proceed with those steps.



Student Packet: 9.2: Redox Reactions: Student Handout

On the AP[®] exam you do not have to balance redox reactions. Know the following typical redox reactions.

Metals in H₂O

Many metals are strong enough reducing agents to reduce water. The reduction potential of water is $2\text{H}_2\text{O} \rightarrow \text{H}_2 + \text{OH}^-$ $E_{\text{red}} = -0.828 \text{ V}$. Any metal with an oxidation potential larger than 0.828 V will reduce H_2O .

Metal in a metal ion solution

If you put a metal into a solution containing a metal ion, go to the table of standard reduction potentials to see if your solid metal will be oxidized and the metal ion reduced. Just look up E_{ox} (the negative of E_{red}) for the metal and E_{red} for the metal ion. If $E_{\text{total}} > 0$, the redox reaction will spontaneously occur.

Metal ion solution mixed with a metal ion solution

Metals can have more than one oxidation state so if you mix solutions containing metal ions, a redox reaction may occur. Go to the table of standard reduction potentials and look at oxidation and reduction half-reactions. If they yield a positive E_{total} , the reaction shall spontaneously occur. Essentially this is as if one metal ion can take electrons from the other metal ion.

Oxidizing Agents

Permanganate ion: MnO_4^-

In acidic solution: $\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + \text{H}_2\text{O}$ $E_{\text{red}} = 1.512 \text{ V}$

In basic solution: $\text{OH}^- + \text{MnO}_4^- \rightarrow \text{MnO}_2 + \text{H}_2\text{O}$

Manganese (II) Oxide: MnO_2

In acidic solution: $\text{MnO}_2 + \text{H}^+ \rightarrow \text{Mn}^{2+} + \text{H}_2\text{O}$ $E_{\text{red}} = 1.229 \text{ V}$

Dichromate ion: $\text{Cr}_2\text{O}_7^{2-}$

In acidic solution: $\text{Cr}_2\text{O}_7^{2-} + \text{H}^+ \rightarrow \text{Cr}^{3+} + \text{H}_2\text{O}$ $E_{\text{red}} = 1.33 \text{ V}$

Hydrogen peroxide: H_2O_2

In acidic solution: $\text{H}_2\text{O}_2 + \text{H}^+ \rightarrow 2\text{H}_2\text{O}$ $E_{\text{red}} = 1.763 \text{ V}$

Know these, then look for a species in solution that can be oxidized.

Metal with an acid

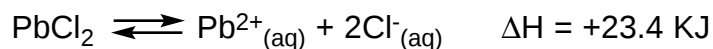
Most metals have positive oxidation potentials. They can be oxidized while H^+ is reduced in an acid solution. Ag and Cu are the exceptions. They have negative oxidation potentials. An additional oxidizing agent is therefore necessary beyond H^+ . In nitric acid, nitrate ion, NO_3^- serves as the oxidizing agent. In sulfuric acid, sulfate ion, SO_4^{2-} serves as the oxidizing agent.

Halogens

The halogens can oxidize/reduce each other. Remember that electronegativity/electron affinity increases as you go up the group. The more electronegative halogen in its diatomic form will oxidize the less electronegative halogen in its ionic form. You can verify this by looking at the half-reaction potentials.

Student Packet: 9.2: Redox Reactions: AssessMe questions

- 1) Predict what effect the addition of HCl has on the position of the equilibrium:



- A) Shift equilibrium to the left.
- B) Shift equilibrium to the right.
- C) No effect.
- D) Cannot be determined.
- E) I'm clueless

- 2) Predict what effect increasing the temperature has on the position of the equilibrium:



- A) Shift equilibrium to the left.
- B) Shift equilibrium to the right.
- C) No effect.
- D) Cannot be determined.
- E) I'm clueless

- 3) Classify H_3PO_4

- A) acid
- B) base
- C) acidic salt
- D) basic salt
- E) I'm clueless

- 4) Classify SrF_2

- A) acid
- B) base
- C) acidic salt
- D) basic salt
- E) I'm clueless

- 5) $2\text{Al} + \frac{3}{2}\text{O}_2 \rightarrow \text{Al}_2\text{O}_3$

What is the oxidation state of Al in the product?

- A) 0
- B) 6
- C) 3
- D) 2
- E) I'm clueless

- 6) $2\text{Cu} + 6\text{H}^+ \rightarrow 2\text{Cu}^{3+} + 3\text{H}_2$

What is the oxidizing agent in this reaction?

- A) Cu
- B) H^+
- C) Cu^{3+}
- D) H_2
- E) I'm clueless

Student Packet: 9.2: Redox Reactions: ShowMe questions

Write the net ionic equation for each of the reactions.

- 1) Solid calcium is added to water.

Manganese (IV) oxide is added to warm, concentrated hydrobromic acid.

An acidified solution of sodium permanganate is added to a solution of potassium sulfite.

- 2) A dilute solution of hydroiodic acid is added to a solution of hydrogen peroxide.

A solution of tin (II) sulfate is added to a solution of iron (III) sulfate.

A strip of copper is immersed in dilute nitric acid.

- 3) Solid sodium is placed in water.

Lead foil is dipped in silver nitrate solution.

A solution of formic acid is mixed with an acidified solution of potassium dichromate.

- 4) Chlorine gas is bubbled through sodium bromide.

Calcium metal is added to a dilute solution of hydrochloric acid.

A solution of tin (II) bromide is added to an acidified solution of potassium permanganate.

- 5) Chlorine gas is bubbled through dilute sodium hydroxide.

Zinc strips are added to a solution of copper (II) sulfate.

A solution of potassium permanganate is mixed with an alkaline solution of lithium sulfite.

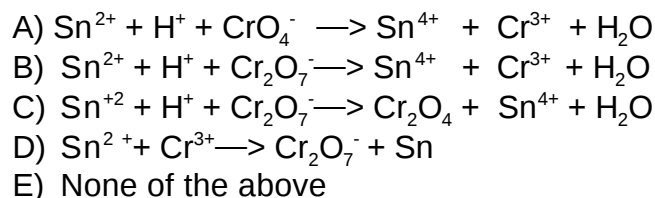
- 6) Powdered iron is added to iron (II) sulfate.

Solid potassium dichromate is added to an acidified solution of potassium iodide.

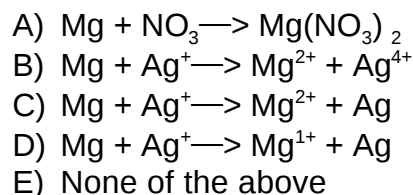
A solution of copper (II) sulfate is electrolyzed using inert electrodes.

Student Packet: 9.2: Redox Reactions: TestMe questions

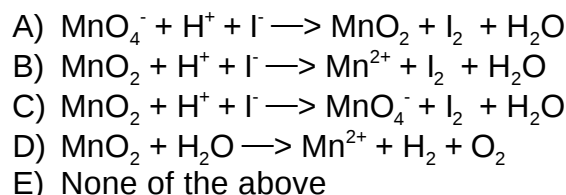
- 1) A solution of tin (II) ions is mixed with an acidic solution of potassium dichromate. Write the net ionic equation. (Don't balance the equation)



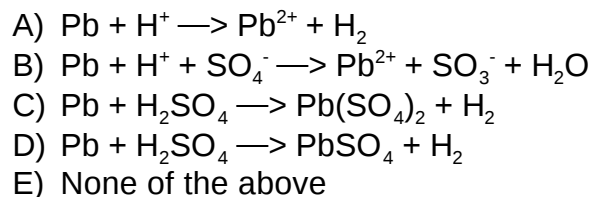
- 2) Generate the net ionic equation for the reaction of solid magnesium when added to a solution of silver nitrate. (Don't balance the equation)



- 3) Concentrated hydroiodic acid solution is added to solid manganese (IV) oxide and the reactants are heated. What is the net ionic equation? (Don't balance the equation)



- 4) Pellets of lead are dropped into hot sulfuric acid. Write the net ionic equation. (Don't balance the equation)



- 5) Copper ribbon is added to 5 M sulfuric acid. Write the net ionic equation. (Don't balance)

