

Redox Reactions

key terms:

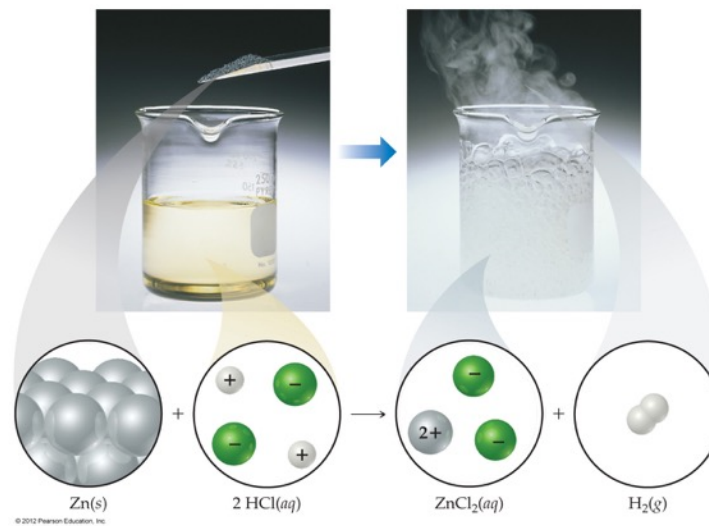
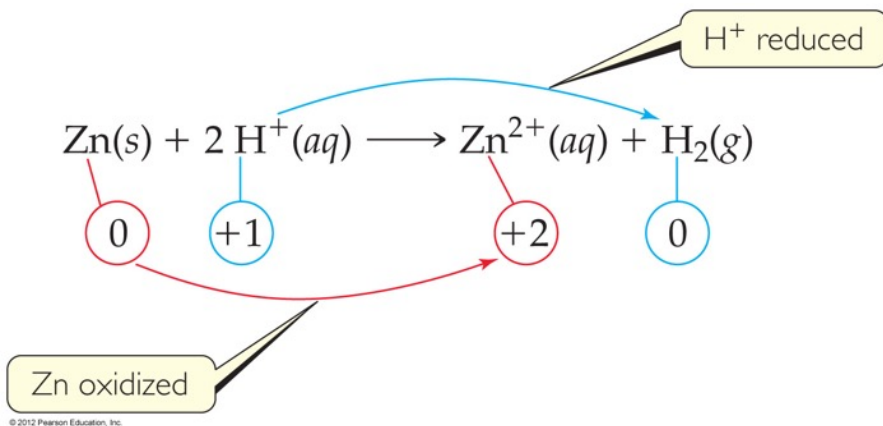
oxidizing agent

reducing agent

oxidation number

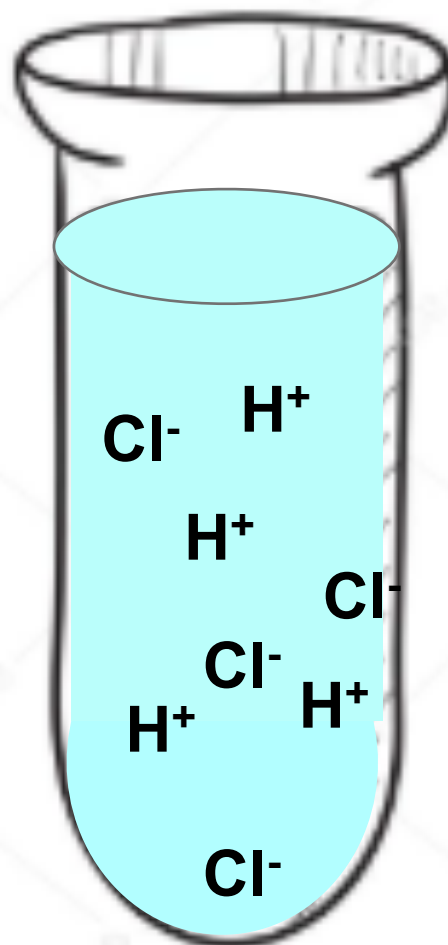
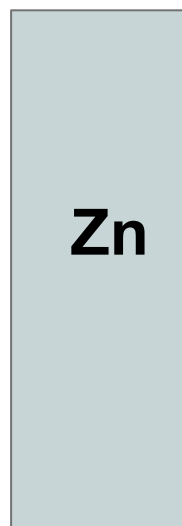
Oxidation Numbers

In order to keep track of what loses electrons and what gains them, we assign **oxidation numbers**.



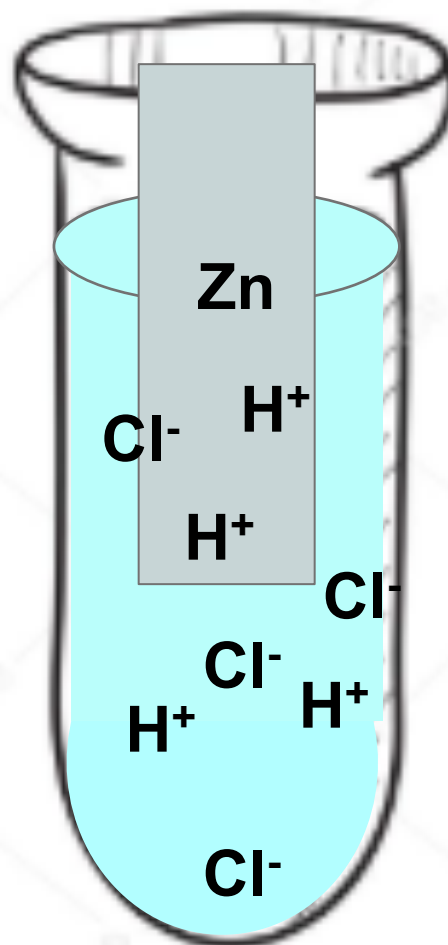
0

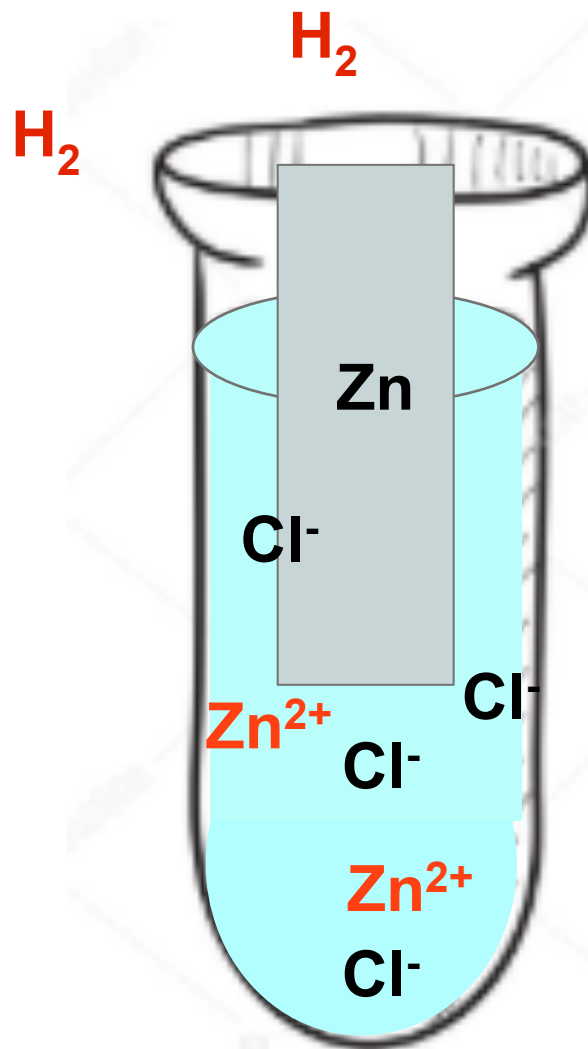
1+



0

1+





Using the activity series provided, write the net ionic chemical equations for the following reaction. If no reaction occurs, simply write NR

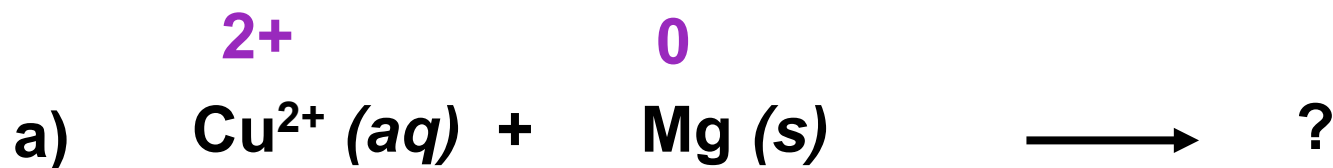


TABLE 4.5 • Activity Series of Metals in Aqueous Solution

Metal		Oxidation Reaction	
Lithium	<i>best reducing agents</i>	$\text{Li}(s) \rightarrow \text{Li}^+(aq) + e^-$	<i>worst oxidizing agents</i>
Potassium		$\text{K}(s) \rightarrow \text{K}^+(aq) + e^-$	
Barium		$\text{Ba}(s) \rightarrow \text{Ba}^{2+}(aq) + 2e^-$	
Calcium		$\text{Ca}(s) \rightarrow \text{Ca}^{2+}(aq) + 2e^-$	
Sodium		$\text{Na}(s) \rightarrow \text{Na}^+(aq) + e^-$	
Magnesium		$\text{Mg}(s) \rightarrow \text{Mg}^{2+}(aq) + 2e^-$	
Aluminum		$\text{Al}(s) \rightarrow \text{Al}^{3+}(aq) + 3e^-$	
Manganese		$\text{Mn}(s) \rightarrow \text{Mn}^{2+}(aq) + 2e^-$	
Zinc		$\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^-$	
Chromium		$\text{Cr}(s) \rightarrow \text{Cr}^{3+}(aq) + 3e^-$	
Iron	$\text{Fe}(s) \rightarrow \text{Fe}^{2+}(aq) + 2e^-$	<i>Ease of oxidation increases</i>	
Cobalt	$\text{Co}(s) \rightarrow \text{Co}^{2+}(aq) + 2e^-$		
Nickel	$\text{Ni}(s) \rightarrow \text{Ni}^{2+}(aq) + 2e^-$		
Tin	$\text{Sn}(s) \rightarrow \text{Sn}^{2+}(aq) + 2e^-$		
Lead	$\text{Pb}(s) \rightarrow \text{Pb}^{2+}(aq) + 2e^-$		
Hydrogen	$\text{H}_2(g) \rightarrow 2\text{H}^+(aq) + 2e^-$		
Copper	$\text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2e^-$		
Silver	$\text{Ag}(s) \rightarrow \text{Ag}^+(aq) + e^-$		
Mercury	$\text{Hg}(l) \rightarrow \text{Hg}^{2+}(aq) + 2e^-$		
Platinum	$\text{Pt}(s) \rightarrow \text{Pt}^{2+}(aq) + 2e^-$		<i>best oxidizing agents</i>
Gold	$\text{Au}(s) \rightarrow \text{Au}^{3+}(aq) + 3e^-$		

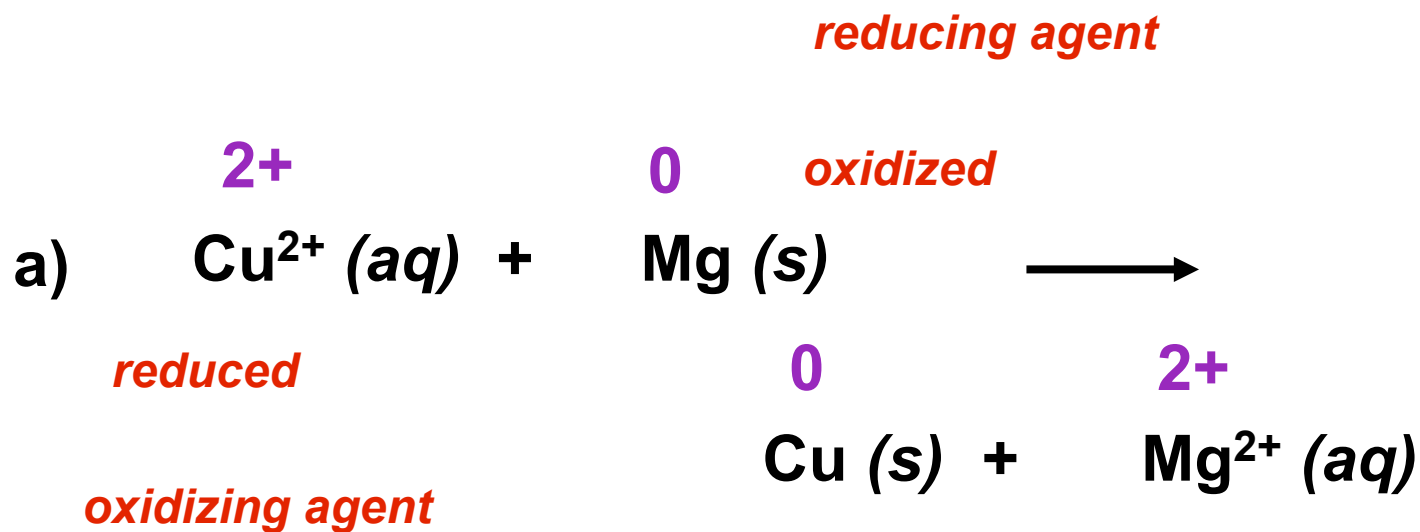
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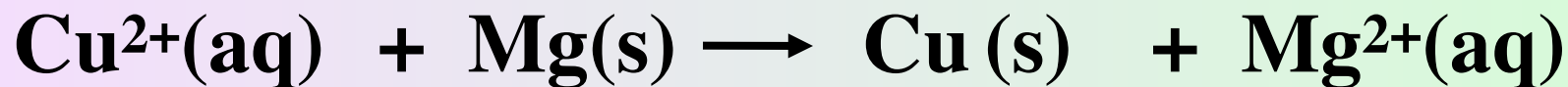
best oxidizing agents

Using the activity series provided, write the net ionic chemical equations for the following reactions. If no reaction occurs, simply write NR



oxidation-reduction reactions

In studying a redox reaction we often think of it as two half reactions.



oxidation



reduction



oxidation-reduction reactions



oxidation



reduction



Balancing Oxidation-Reduction Equations

Using the activity series provided, write the net ionic chemical equations for the following reactions. If no reaction occurs, simply write NR

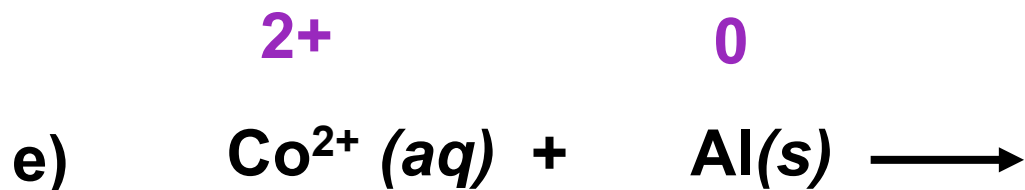


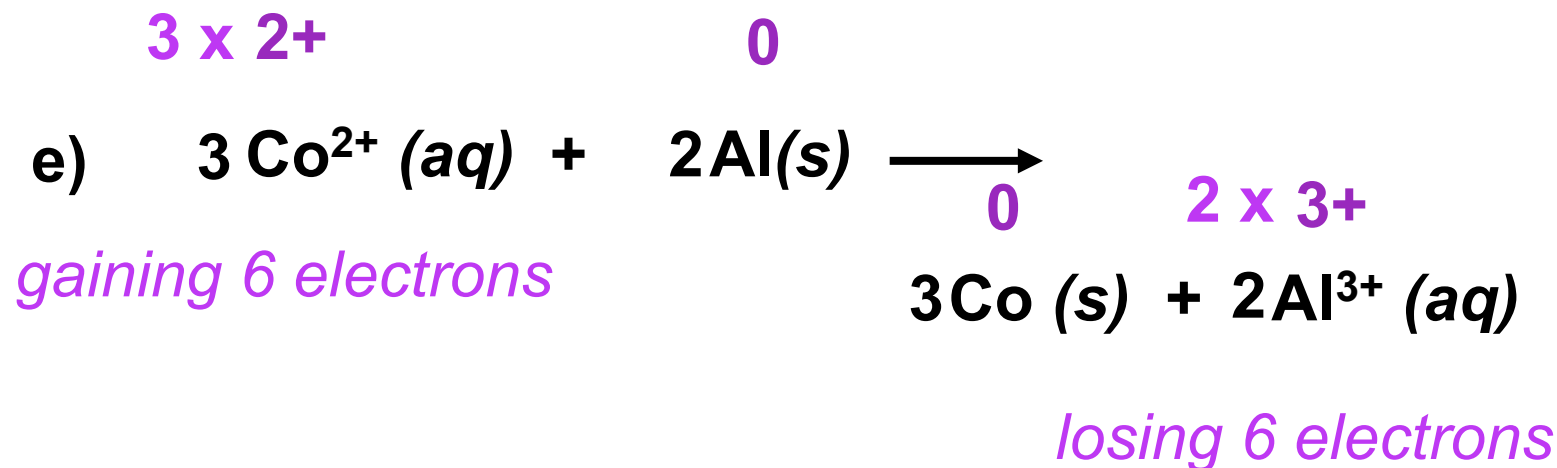
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best oxidizing agents

Using the activity series provided, write the net ionic chemical equations for the following reactions. If no reaction occurs, simply write NR



Half-Reaction Method in Acid

- 1. Write the unbalanced equation in ionic form.**
- 2. Separate the equation into two half-reactions.**
- 3. Balance each half reaction (except for O and H).**
- 4. In acid solution, balance O by adding H_2O and H by adding H^+**
- 5. Balance the charges by adding electrons.**
- 6. Add the half reactions**
- 7. Check to make sure atoms and charges are balanced**

Example

Balance the following equation for the reaction in acid solution. *



***All species are (aq)**

1. Write the unbalanced equation in ionic form.



2. Separate the equation into two half-reactions



3. Balance each half reaction (except for O and H).

4. In acid solution, balance O by adding H₂O and H by adding H⁺



5. Balance the charges by adding electrons.

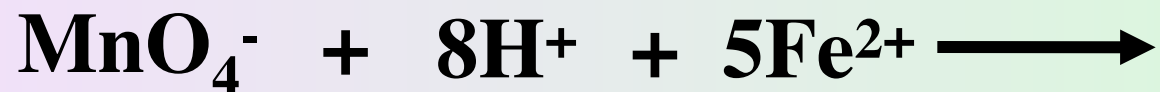
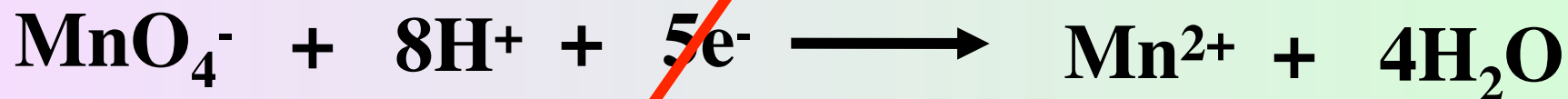
You need the same number of electrons on both sides of the equation.



5. Balance the charges by adding electrons.



6. Add the half reactions



This is the balanced equation

Half-Reaction Method in Base

- 1. Use the half-reaction method as specified for acidic solutions to obtain the final balanced equation *as if H^+ ions were present.***
- 2. Add the number of OH^- ions to both sides of the equation to turn the remaining H^+ ions to H_2O**
- 3. Eliminate waters that appear on both sides of the equation.**

Example

Balance the following equation for the reaction in basic solution. *



*All species are (aq)

Example

Balance the following equation for the reaction in basic solution. *



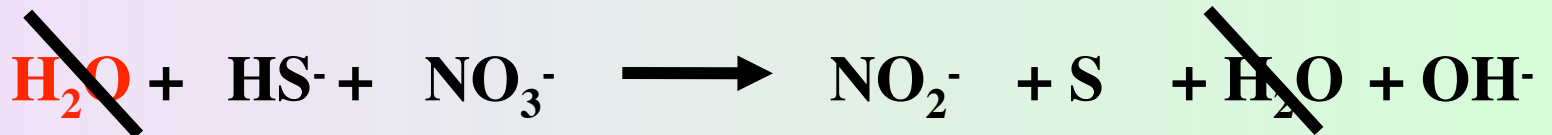
Example

Balance the following equation for the reaction in basic solution. *



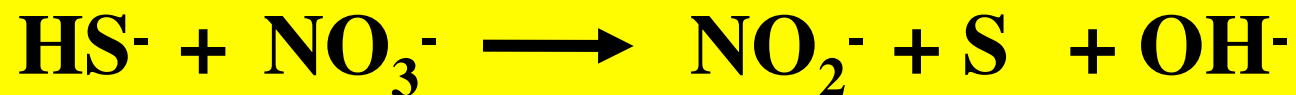
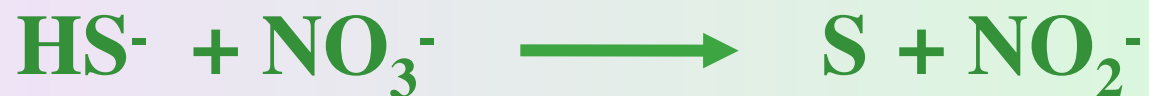
Example

Balance the following equation for the reaction in basic solution. *



Example

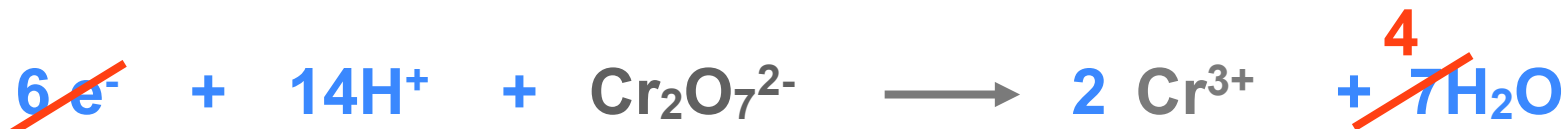
Balance the following equation for the reaction in basic solution. *



Complete and balance the following equations,
and identify the oxidizing and reducing agents:

acidic solution

a)



Complete and balance the following equations, and identify the oxidizing and reducing agents:

c)

acidic solution



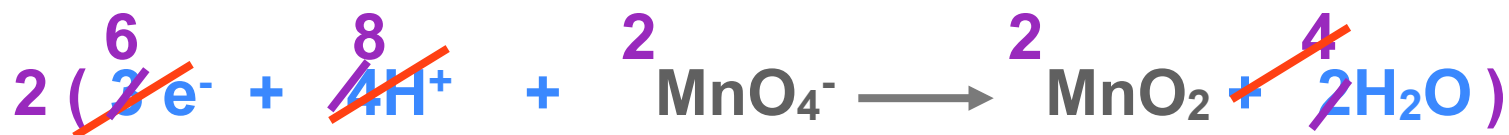
Half-Reaction Method in Base

- 1. Use the half-reaction method as specified for acidic solutions to obtain the final balanced equation *as if H^+ ions were present*.**
- 2. Add the number of OH^- ions to both sides of the equation to turn the remaining H^+ ions to H_2O**
- 3. Eliminate waters that appear on both sides of the equation.**

Complete and balance the following equations, and identify the oxidizing and reducing agents:

d)

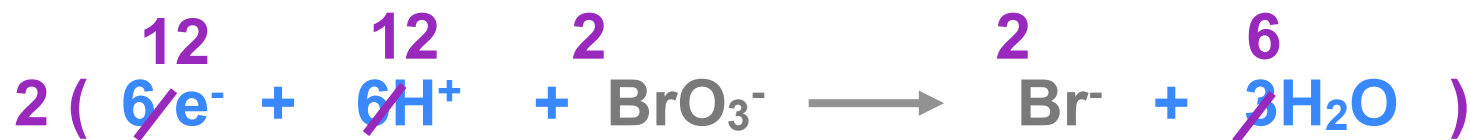
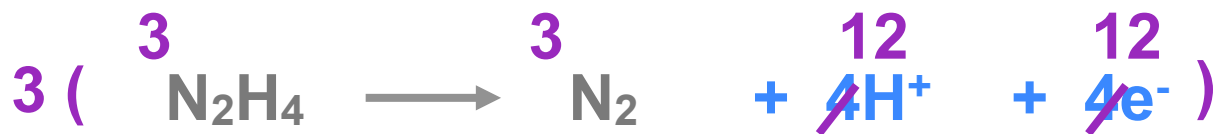
basic solution



Complete and balance the following equations, and identify the oxidizing and reducing agents:

e)

acidic solution



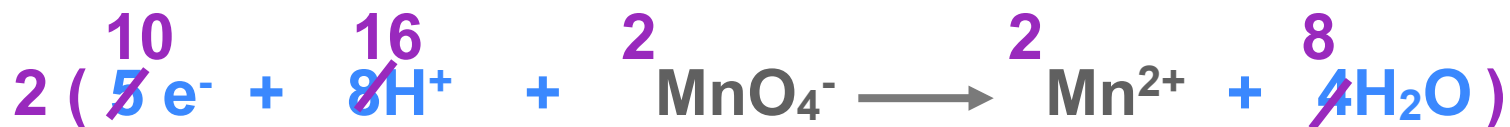
Complete and balance the following equations,
and identify the oxidizing and reducing agents:

f) *basic solution*



Complete and balance the following equations, and identify the oxidizing and reducing agents:

acid solution



Half-Reaction Method in Acid

1. Write the unbalanced equation in ionic form.
2. Separate the equation into two half-reactions.
3. Balance each half reaction (except for O and H).
4. In acid solution, balance O by adding H_2O
5. In acid solution, balance H by adding H^+
6. Balance the charge in the half reaction by adding electrons.
7. multiply by half reactions by an integer to make the electrons lost equal to the electrons gained
8. Add the half reactions (canceling anything that is redundant)
- 9 Check to make sure atoms and charges are balanced

Half-Reaction Method in base

1. Use the acid method, then add the number of OH^- ions to both sides of the equation to turn the remaining H^+ ions to H_2O
2. Eliminate waters that appear on both sides of the equation.

problems 34-37 refer to the chemical reaction below



which atom is acting as the oxidizing agent?

What is the oxidation state of zinc ion at the end of the reaction?

Which element is reduced during the reaction reaction?

Complete and balance the following equations, and identify the oxidizing and reducing agents:

acidic solution

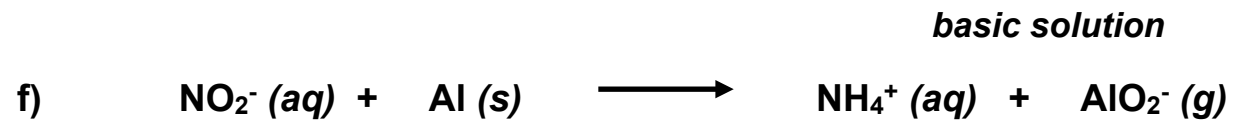


acidic solution



acidic solution





Example

Balance the following equation for the reaction in acid solution. *



1. Write the unbalanced equation in ionic form.



2. Separate the equation into two half-reactions

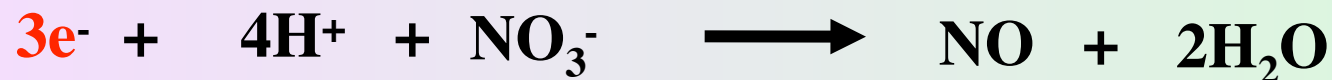


3. Balance each half reaction (except for O and H).

4. In acid solution, balance O by adding H₂O and H by adding H⁺



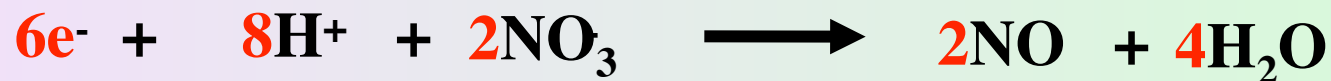
5. Balance the charges by adding electrons.



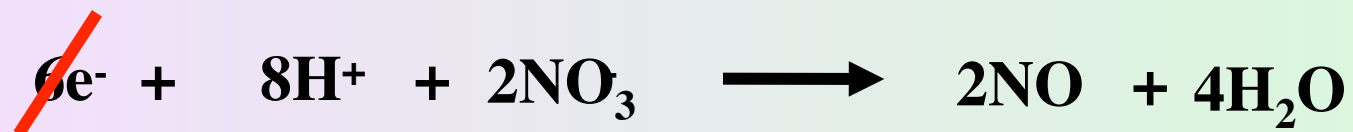
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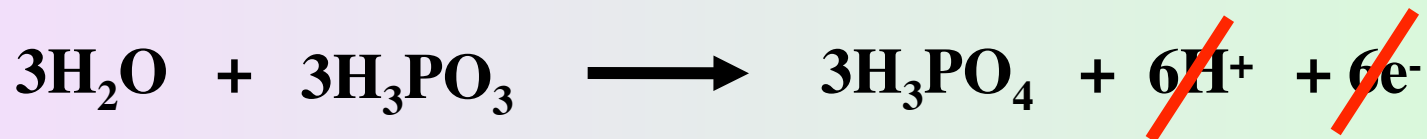
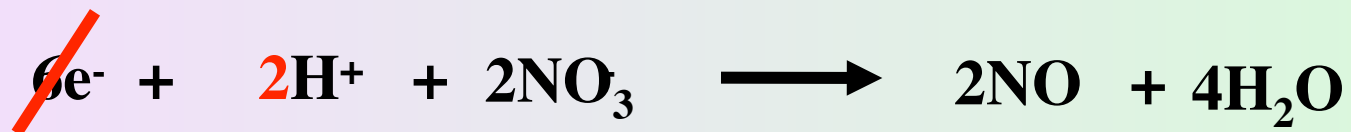
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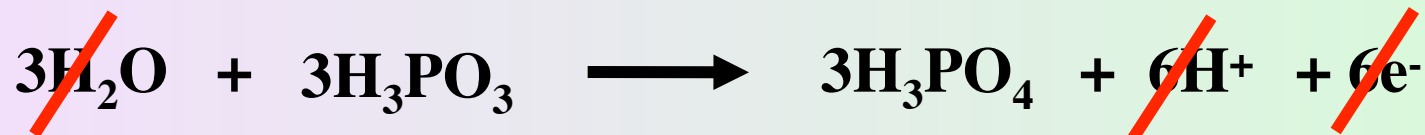
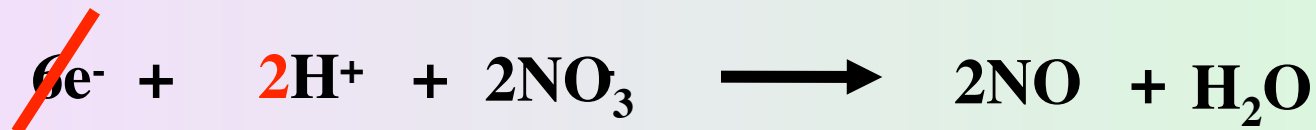
6. Add the half reactions



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This is the balanced equation

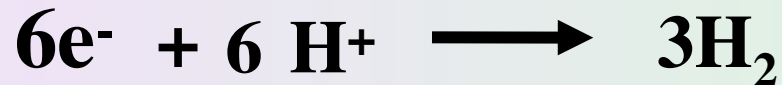
Example

Balance the following equation for the reaction in acid solution. *

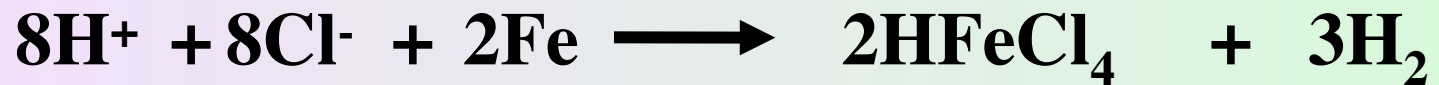
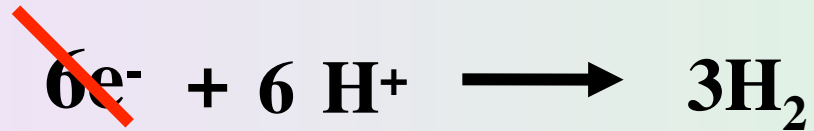


Example

Balance the following equation for the reaction in acid solution. *



Example



or

