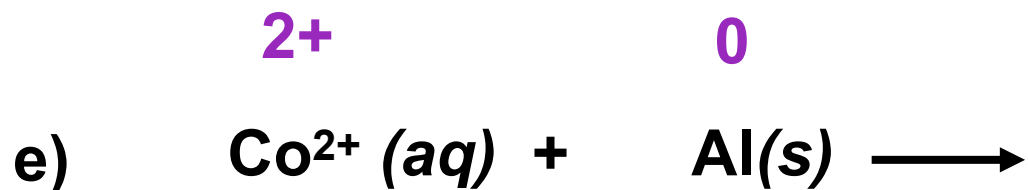


# **Balancing Oxidation-Reduction Equations**

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



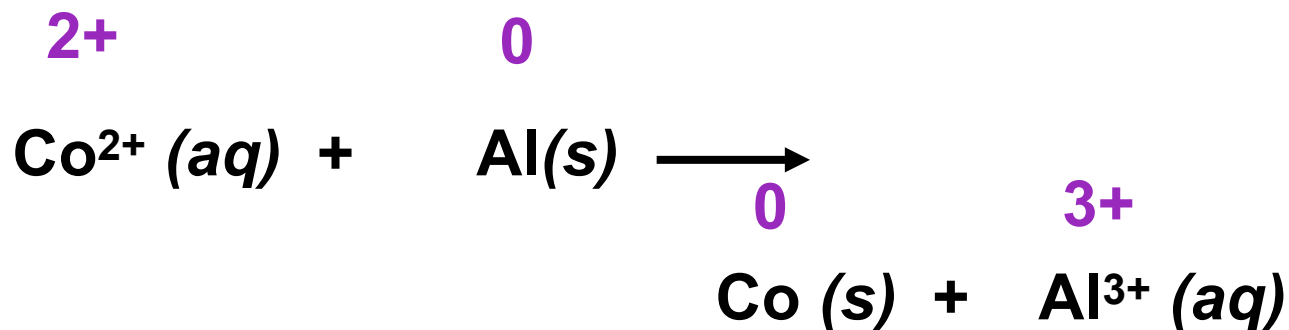
**TABLE 4.5 • Activity Series of Metals in Aqueous Solution**

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



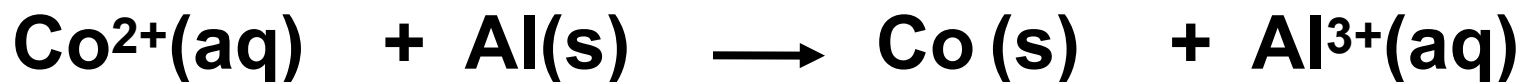
*best oxidizing agents*

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

the number of electrons lost has to be equal to the number of electrons gained.



**oxidation**

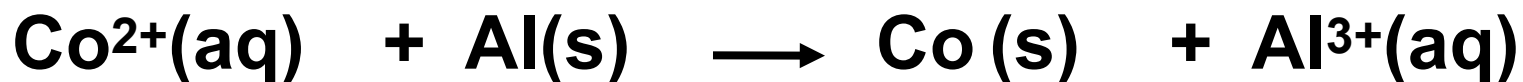


**reduction**



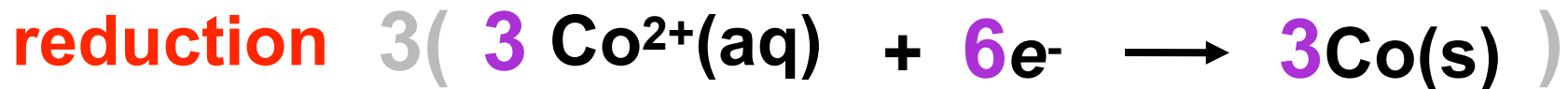
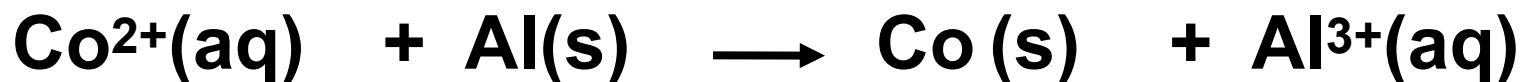
# oxidation-reduction reactions

the number of electrons lost has to be equal to the number of electrons gained.



# oxidation-reduction reactions

the number of electrons lost has to be equal to the number of electrons gained.



# oxidation-reduction reactions

Adding the half reactions to gather gives us the overall redox reaction.

**oxidation**



**reduction**

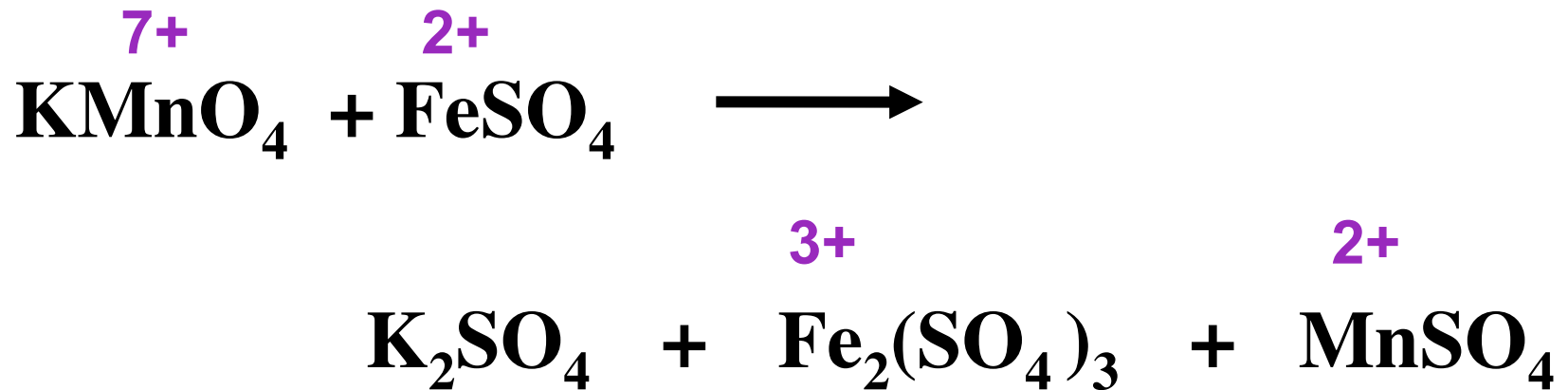




# Example

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is this a redox reaction?



\*All species are (aq)

# Example

---

Balance the following equation



\*All species are (aq)

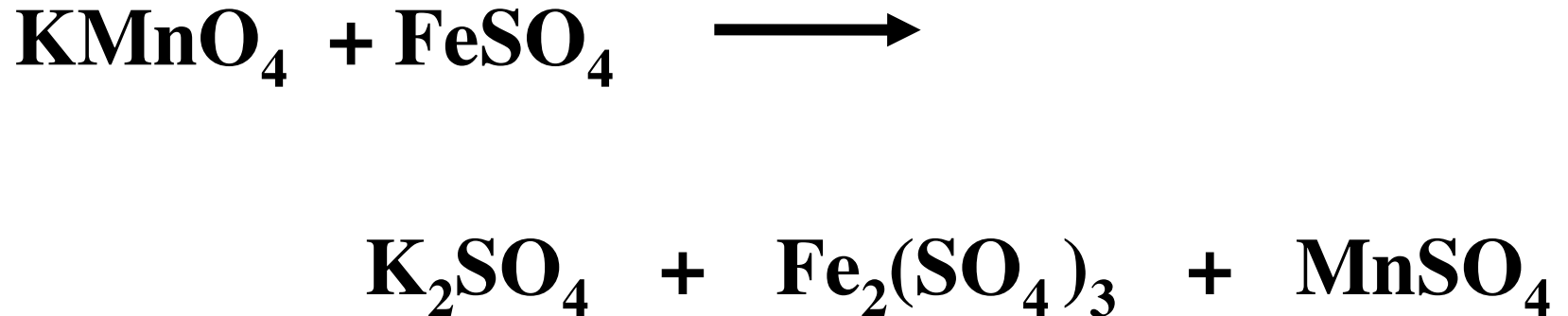
## **Half-Reaction Method in Acid**

- 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  $\text{H}_2\text{O}$  and H by adding  $\text{H}^+$**
- 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)

# 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 H<sub>2</sub>O and H by adding H<sup>+</sup>**



## 5. Balance the charges by adding electrons.

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

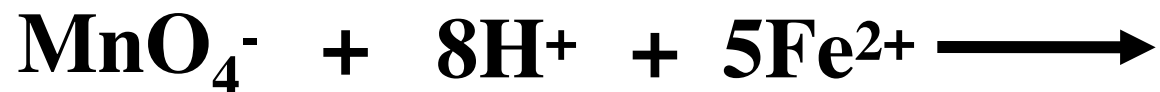
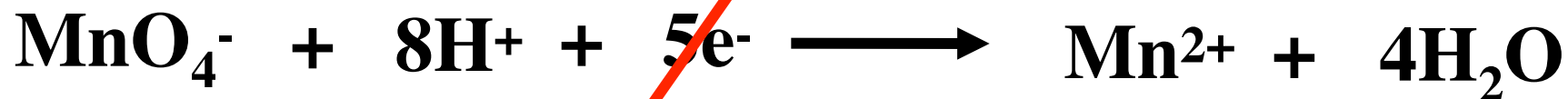




**5. Balance the charges by adding electrons.**



## 6. Add the half reactions



**This is the balanced equation**

# Example

---

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



**\*All species are (aq)**

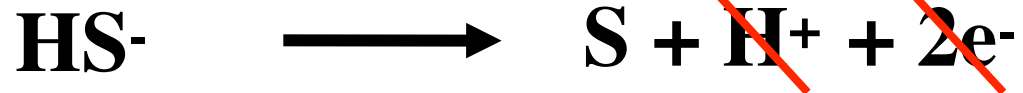
## Half-Reaction Method in Base

- 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

---

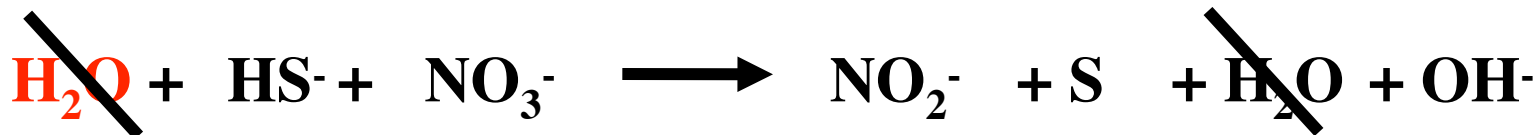
Balance the following equation for the reaction in basic solution. \*



# Example

---

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

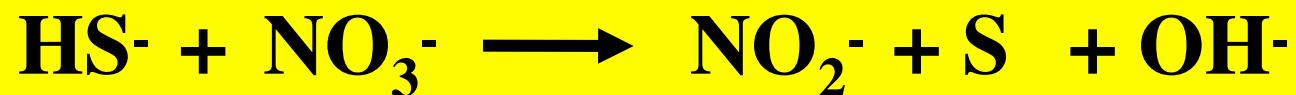




# Example

---

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*

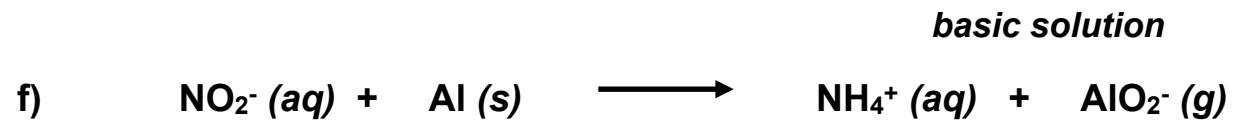


*acidic solution*



*acidic 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  $\text{H}_2\text{O}$
5. In acid solution, balance H by adding  $\text{H}^+$
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

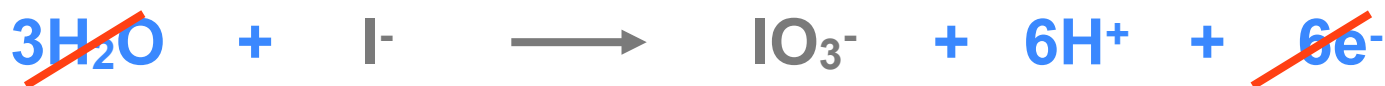
## Half-Reaction Method in base

1. Use the acid method, then add the number of  $\text{OH}^-$  ions to both sides of the equation to turn the remaining  $\text{H}^+$  ions to  $\text{H}_2\text{O}$
2. Eliminate waters that appear on both sides of the equation.

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

*acidic solution*

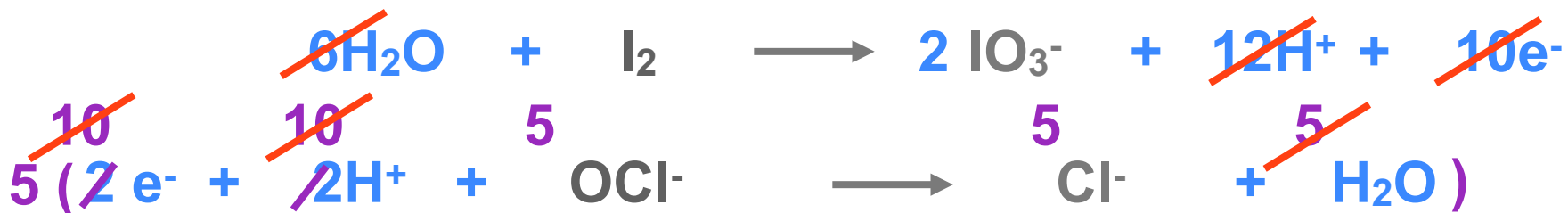
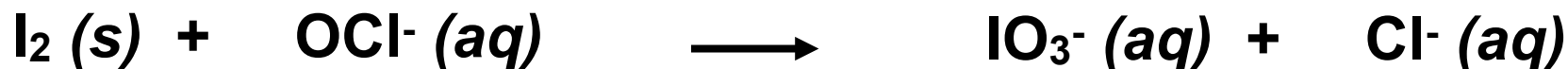
a)



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

*acidic solution*

b)



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

c)

*acidic solution*

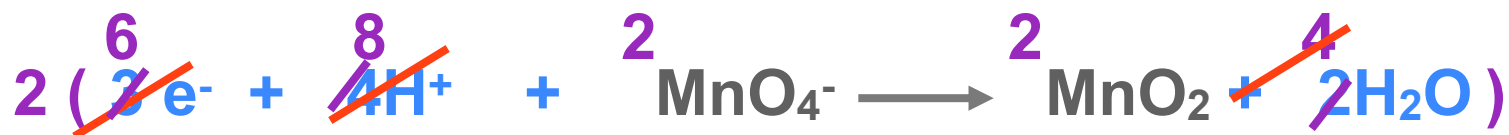




Complete and balance the following equations, and identify the oxidizing and reducing agents:

d)

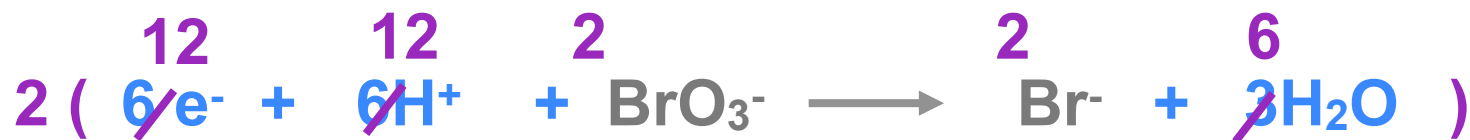
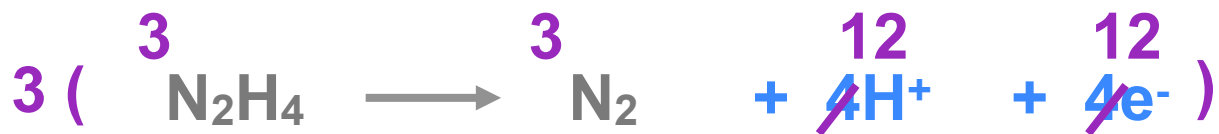
*basic solution*



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

e)

*acidic solution*



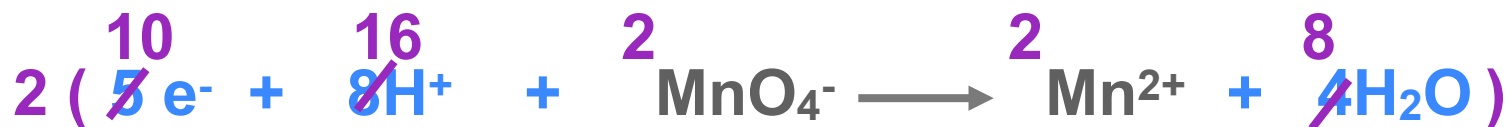
Complete and balance the following equations,  
and identify the oxidizing and reducing agents:

f) *basic solution*



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

*acid solution*



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?

# Example

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**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**



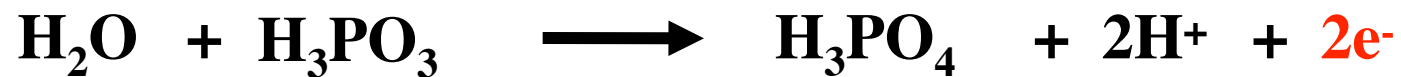
**3. Balance each half reaction (except for O and H).**

**4. In acid solution, balance O by adding H<sub>2</sub>O and H by adding H<sup>+</sup>**

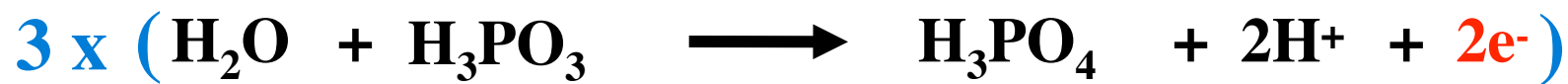
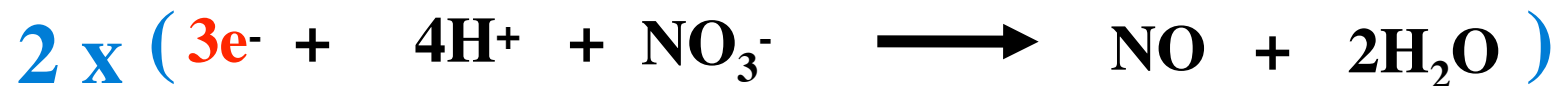




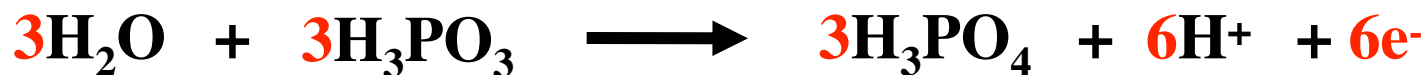
## 5. Balance the charges by adding electrons.



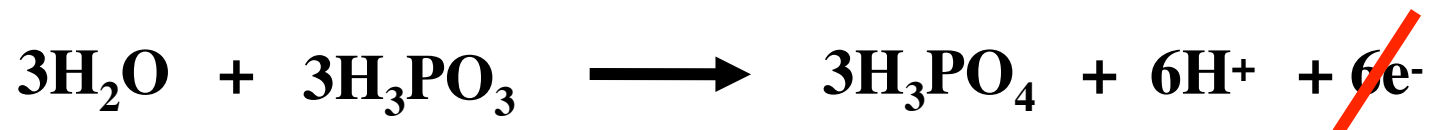
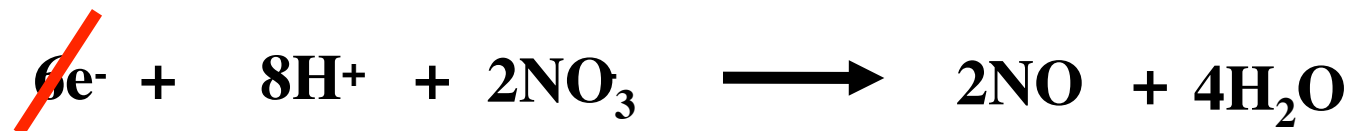
## 5. Balance the charges by adding electrons.



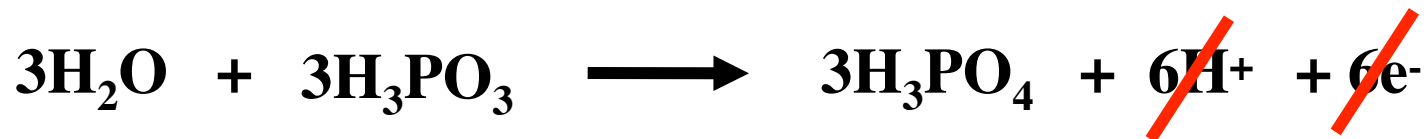
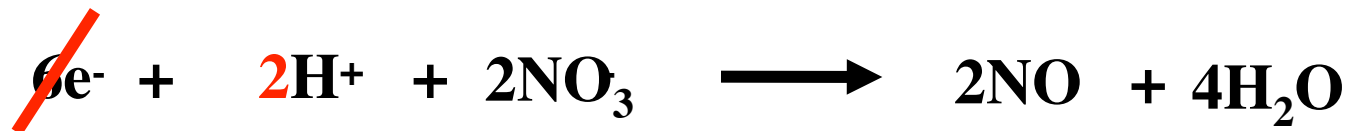
## 5. Balance the charges by adding electrons.



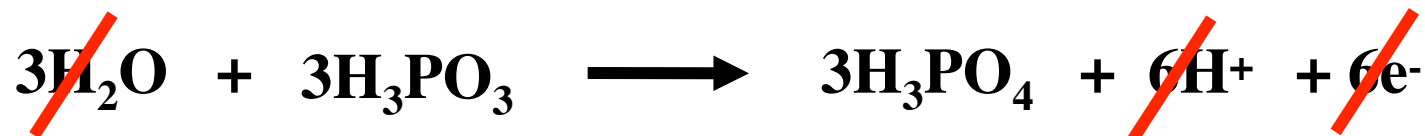
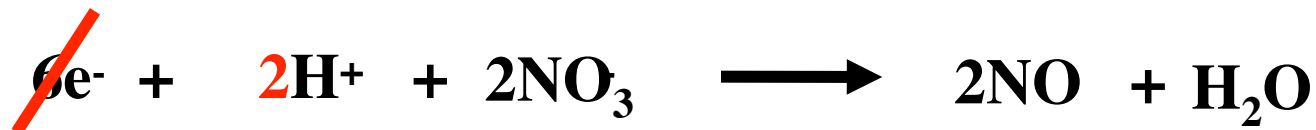
## 6. Add the half reactions



## 6. Add the half reactions



## 6. Add the half reactions



## 6. Add the half reactions

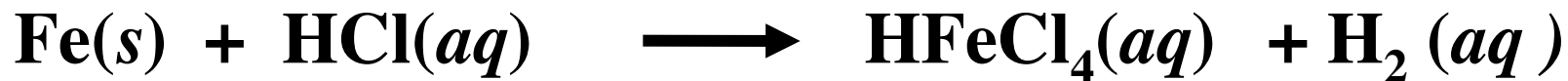


**This is the balanced equation**

# Example

---

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

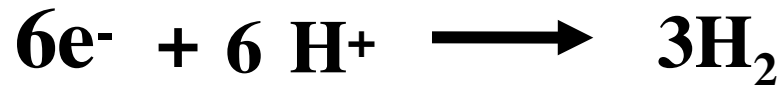
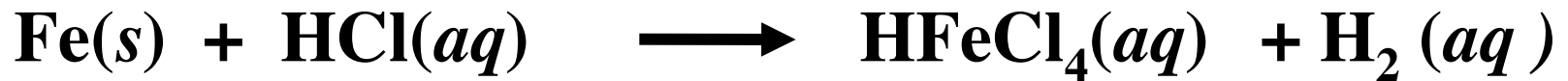




# Example

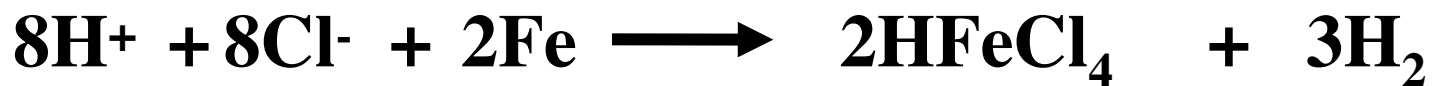
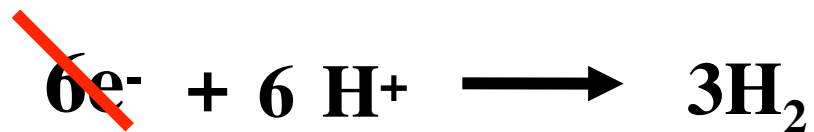
---

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



# Example

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or



