

# Chapter 17 Part 2

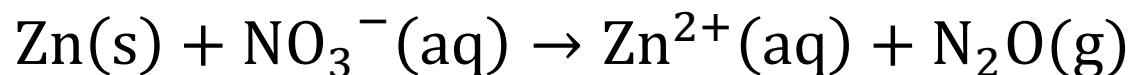
Dr. Turner

# Balancing Redox Reactions in Acidic Solution

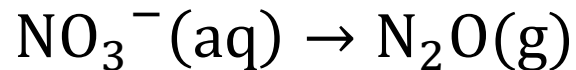
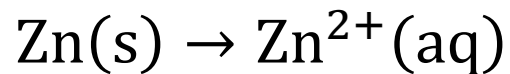
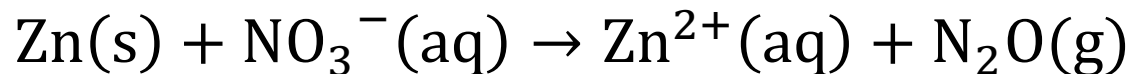
- Write the equations for the oxidation and reduction half-reactions
- In each half-equation
  - ▣ Balance the atoms of all the elements except O and H
  - ▣ Balance oxygen using  $\text{H}_2\text{O}$
  - ▣ Balance hydrogen using  $\text{H}^+$
  - ▣ Balance charge using electrons
- If necessary, equalize the number of electrons in the oxidation and reduction half equations by multiplying one or both half-equations by appropriate integers
- Add the half-equations, then cancel species common to both sides of the over-all equation
- Check that numbers of atoms and charges balance

# Balancing Redox Reactions in Acidic Solution

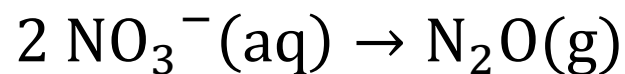
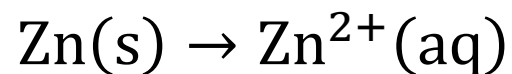
Write the balance equation for the reaction below in acidic solution.



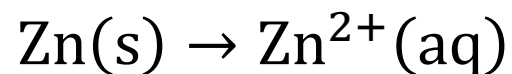
# Write Half Reactions



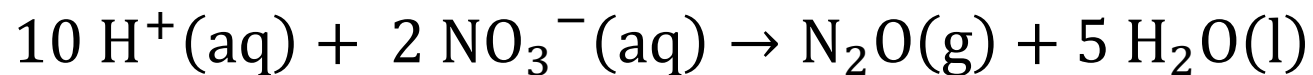
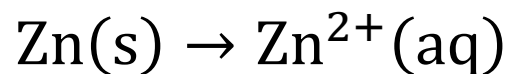
Balance the atoms of all the elements except O and H



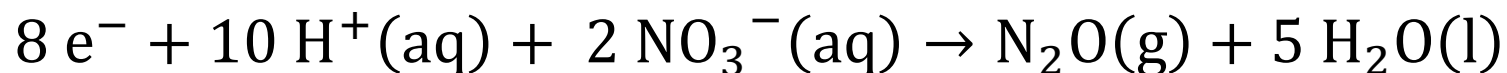
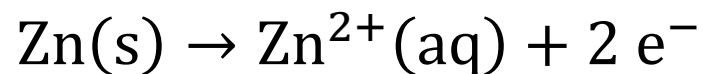
# Balancing O using H<sub>2</sub>O



# Balancing H using $\text{H}^+$

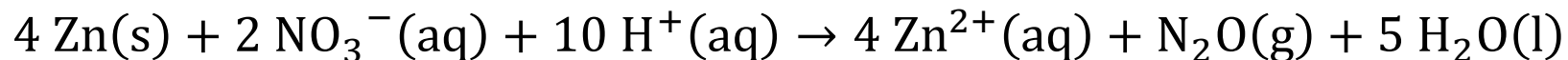
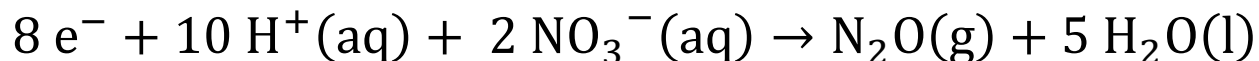
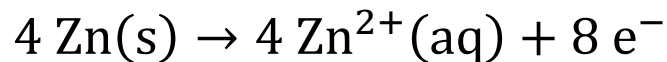
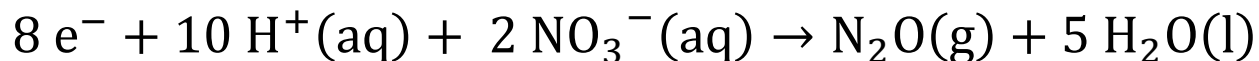
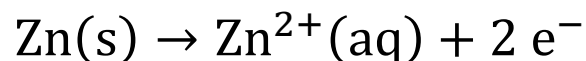


# Balancing Charge by Adding Electrons



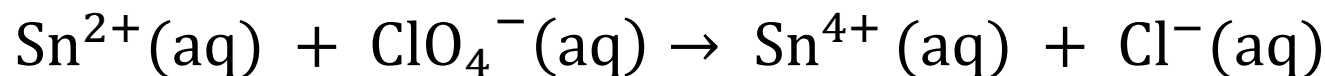


# Combine the Half-Equations



# Balancing redox reactions in acidic sol'n

Balance the following reaction in acidic solution.

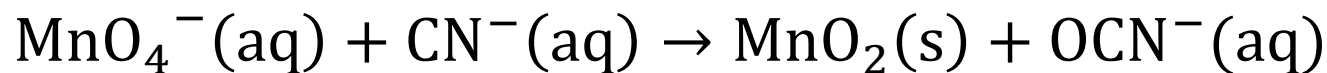


# Balancing Redox Reactions in Basic Solution

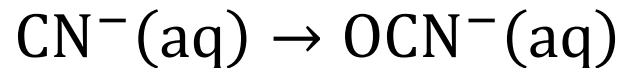
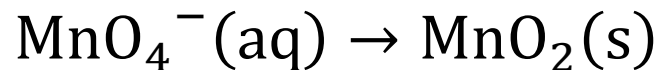
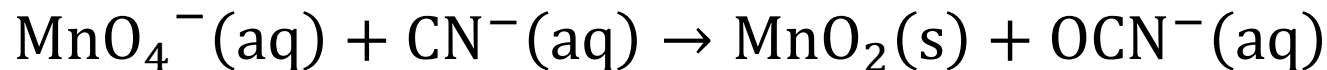
- Balance the equation as if the reaction were occurring in acidic medium
- To both sides of the overall equation obtained, add a number of  $\text{OH}^-$  that is equal to the number of  $\text{H}^+$  ions.
- On the side of the overall equation containing both  $\text{H}^+$  and  $\text{OH}^-$  ions, combine them to form  $\text{H}_2\text{O}$  molecules.
- If  $\text{H}_2\text{O}$  molecules now appear on both sides of the overall equation, cancel the same number from each side, leaving a remainder of  $\text{H}_2\text{O}$  on just one side.
- Check that numbers of atoms and charges balance.

# Balancing Redox Reactions in Basic Solution

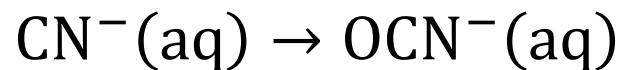
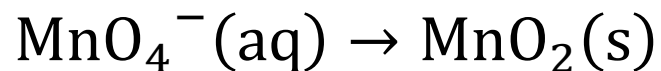
Balance the equation for the reaction in which permanganate ion oxidizes cyanide ion to cyanate ion in basic solution and is itself reduced to  $\text{MnO}_2(\text{s})$ .



# Write Half Reactions

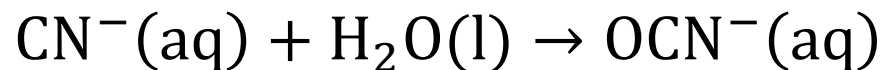


Balance the atoms of all the elements except O and H

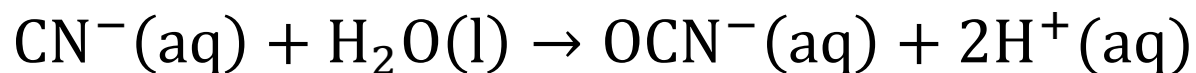
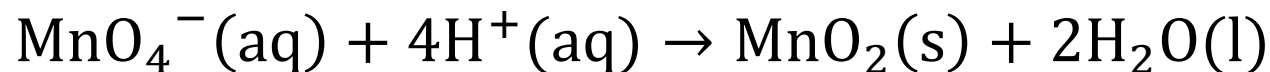


# Balancing O using H<sub>2</sub>O

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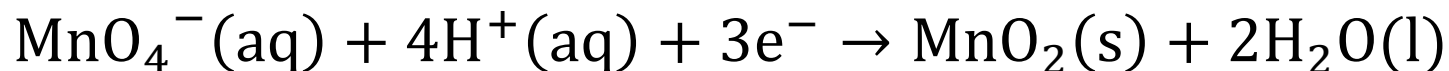


# Balancing H using $\text{H}^+$

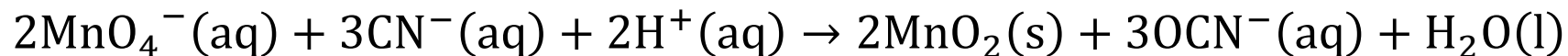
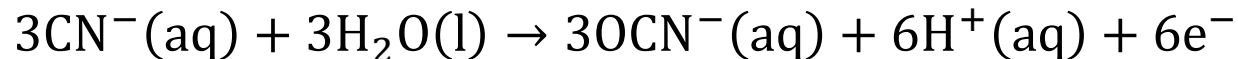
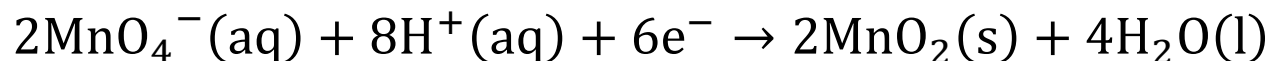
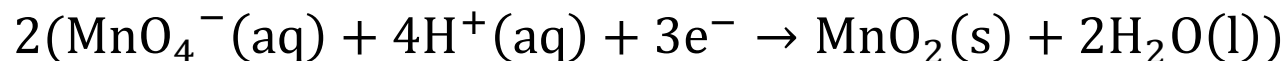




# Balancing Charge by Adding Electrons

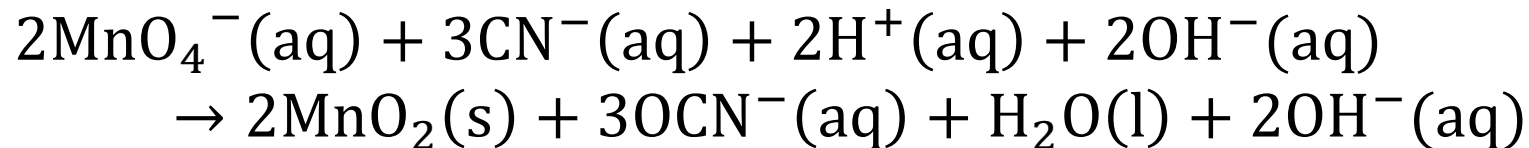


# Combine the Half-Equations



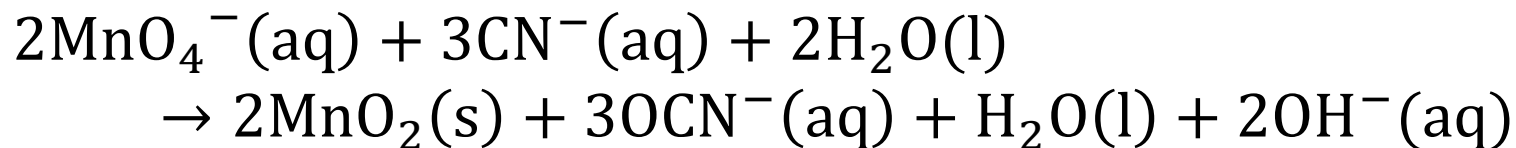
# Add $\text{OH}^-$ to both sides for every $\text{H}^+$

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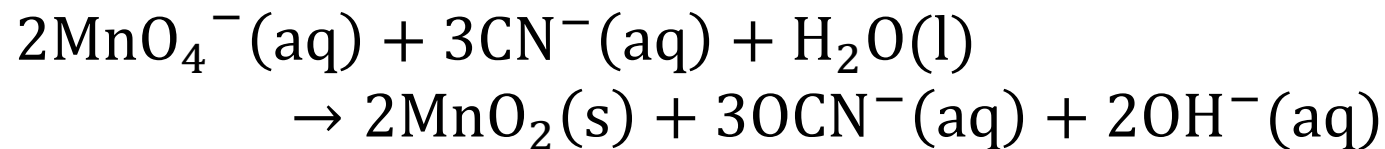


# Combine $\text{H}^+$ with $\text{OH}^-$ to make $\text{H}_2\text{O}$

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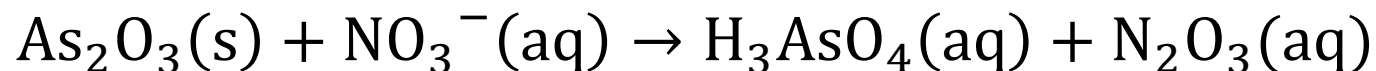


# Cancel extra H<sub>2</sub>O



# Balancing redox reactions in basic sol'n

Balance the following reaction in basic solution.



# Half Reactions

Identify the reduction half-reaction for:



- A.  $2 \text{Na(s)} + 2 \text{e}^- \rightarrow 2 \text{Na}^+(\text{aq})$
- B.  $\text{F}_2(\text{g}) + 2 \text{e}^- \rightarrow 2 \text{F}^-(\text{aq})$
- C.  $2 \text{Na(s)} \rightarrow 2 \text{Na}^+(\text{aq}) + 2 \text{e}^-$
- D.  $\text{F}_2(\text{g}) \rightarrow 2 \text{F}^-(\text{aq}) + 2 \text{e}^-$

# Balancing Redox Reactions

Write the balanced overall redox equation for the half-reactions below.



- A.  $\text{Fe}^{3+}(\text{aq}) + \text{Br}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{s}) + \text{Br}_2(\text{l})$
- B.  $\text{Fe}^{3+}(\text{aq}) + 2 \text{Br}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{s}) + \text{Br}_2(\text{l})$
- C.  $2 \text{Fe}^{3+}(\text{aq}) + 4 \text{Br}^{-}(\text{aq}) \rightarrow 2 \text{Fe}(\text{s}) + 2 \text{Br}_2(\text{l})$
- D.  $2 \text{Fe}^{3+}(\text{aq}) + 6 \text{Br}^{-}(\text{aq}) \rightarrow 2 \text{Fe}(\text{s}) + 3 \text{Br}_2(\text{l})$