

## Balancing Oxidation-Reduction Equations

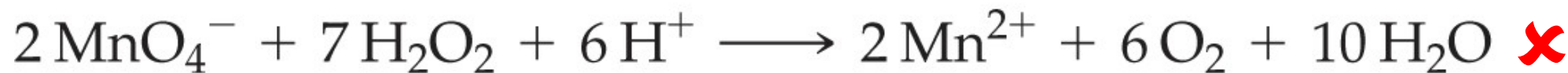
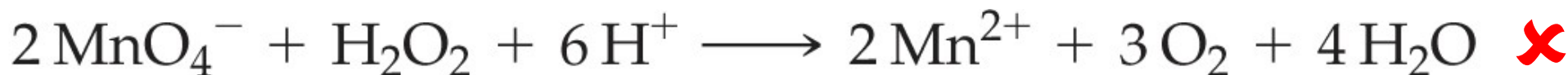
We can't count on "balancing by inspection" or even the more general procedure we learned earlier.

We need to use a method specialized to redox reactions.

Especially because we will be expected to bring in  $\text{H}^+$ ,  $\text{OH}^-$ , and  $\text{H}_2\text{O}$  into the reaction equation even if they are not initially in the given equation, as we will normally consider redox reactions in aqueous solution.

## Balancing Oxidation-Reduction Equations

The following equations are all balanced in terms of atoms and charges, but only one is balanced in terms of the number of electrons lost and gained.



We need a method that exposes the electrons lost in oxidation and gained in reduction. Then we can make sure they are equal.

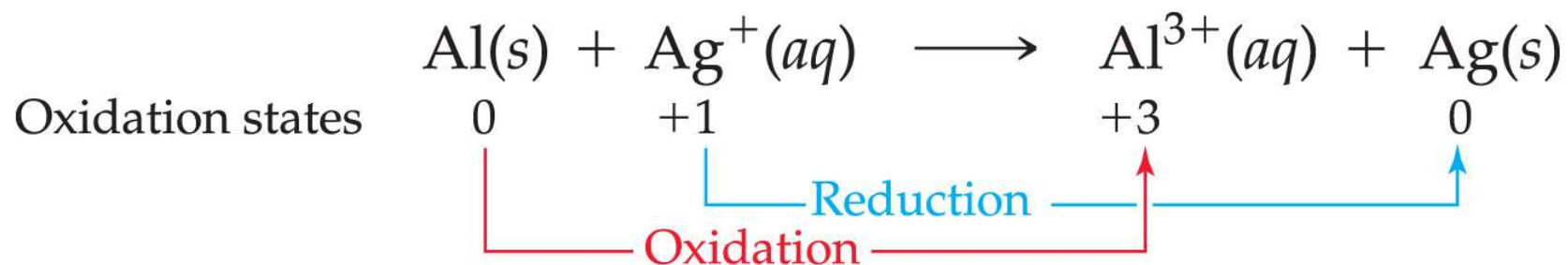
**Half-reaction method does that.**

## A General Procedure for Balancing Redox Reactions Using the Half-Reaction Method

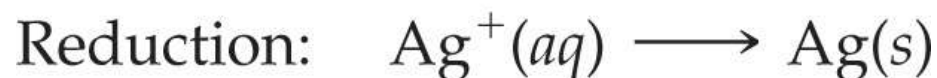
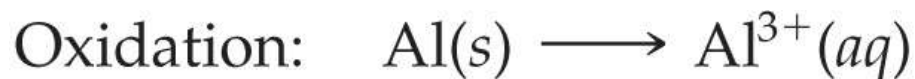
1. Assign oxidation states.
2. Separate the overall reaction into two half-reactions: oxidation and reduction
3. Balance each half-reaction with respect to number of atoms of each element.
4. Balance each half-reaction with respect to charge by adding electrons.
5. Make the number of electrons lost/gained in both half-reactions equal by multiplying by appropriate factors
6. Add the two half-reactions together.
7. Verify that the reaction is balanced.

## Balancing Redox Equations

- Assign oxidation numbers to all atoms to determine what is being oxidized and what is being reduced.



- Divide the reaction into two half-reactions, one for oxidation and one for reduction.



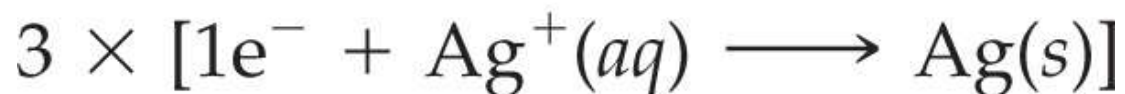
## Balancing Redox Equations

- In this case, the half-reactions are already balanced with respect to mass—the number of each type of atom on both sides of each half-reaction is the same.
- However, the equations are not balanced with respect to charge.
- We balance the charge of each half-reaction individually by adding the appropriate number of electrons to make the charges on both sides equal.



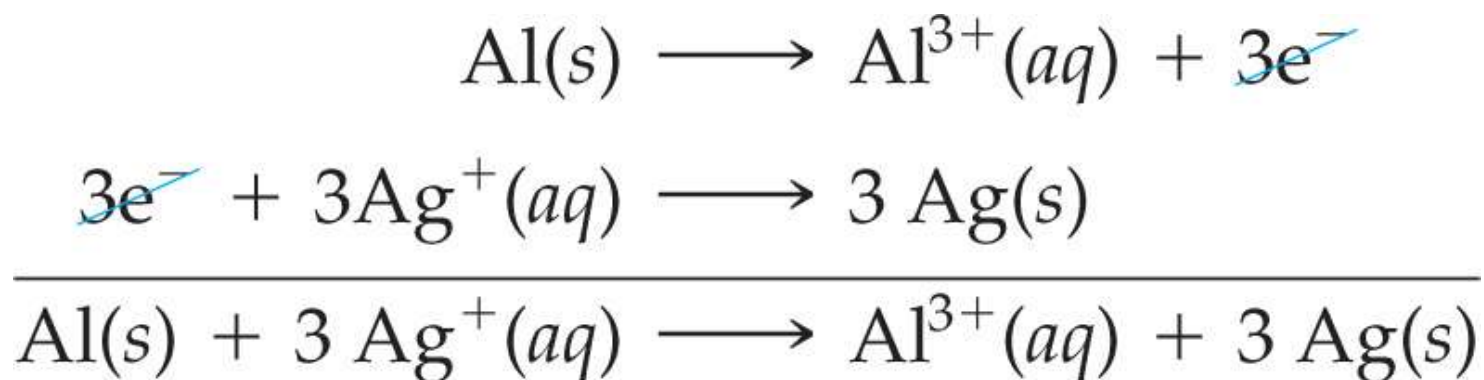
## Balancing Redox Equations

- Since these half-reactions must occur together, the number of electrons lost in the oxidation half-reaction must equal the number gained in the reduction half-reaction.
- We equalize these by multiplying one or both half-reactions by appropriate whole numbers to equalize the electrons lost and gained.



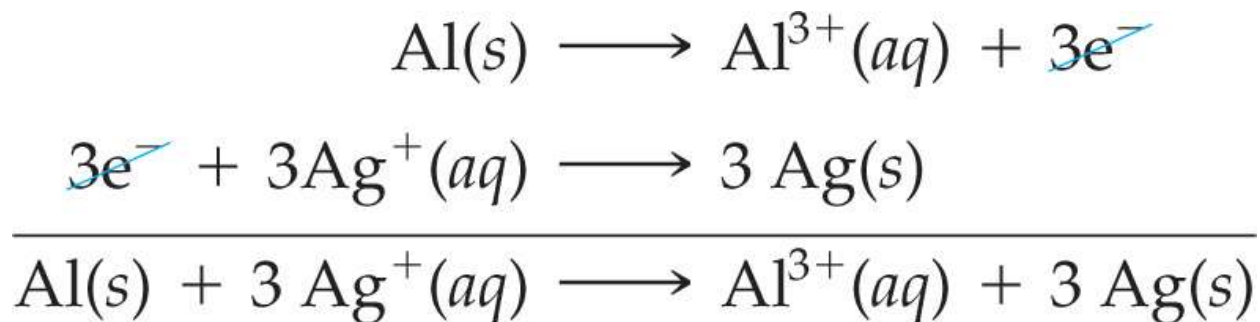
## Balancing Redox Equations

- Add the half-reactions together, canceling electrons and other species as necessary.



## Balancing Redox Equations

- Verify that the equation is balanced, with respect to both mass and charge.



| Reactants | Products  |
|-----------|-----------|
| 1 Al      | 1 Al      |
| 3 Ag      | 3 Ag      |
| +3 charge | +3 charge |

- Notice that the charge need not be zero on both sides of the equation—it just has to be *equal* on both sides. The equation is balanced.

## Balancing Redox Reactions in Aqueous Solutions

Balancing redox reactions in aqueous solution might look like alchemy where the initial equation might be missing an element or two (or more) on one side of the equation or the other!

You might feel like you are doing things we are not supposed to do in balancing, such as adding new reactants or products to the equation. That's because aqueous-solution redox reactions are often given in “bare-bones” form.

- We are expected to “fill in the blanks” and create a complete equation while balancing it.
- Actually, it's easier to start balancing without the species that are not involved in redox and then add them later. We just need to get used to it.

# Balancing Redox Reactions in Aqueous Solutions

## -- “ionic” or “net ionic” equations

The “bare-bones” (i.e. skeletal) reactions are usually given in “net ionic” form

- Some authors seem to use the term “ionic equation” to mean “net ionic equation”; so beware!

“ionic equation” and “net ionic equation” are analogous to those for double-displacement (metathesis) reactions:

- Ionic equation: include the “counter-ions” that would allow you to combine ions into neutral chemical formulas
- Net-ionic equation: do not include spectator ions (those “counter-ions” of ions that are actually involved in redox

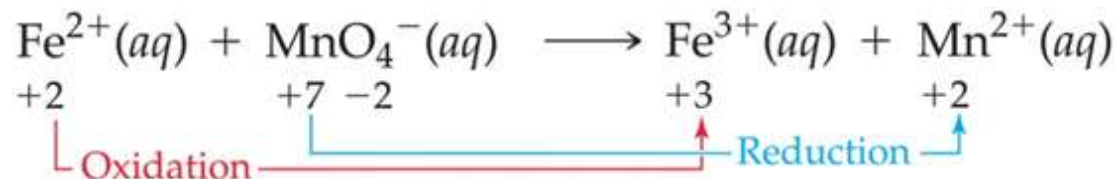
## Balancing Redox Reactions in Acidic Conditions

- Write both half reactions (oxidation & reduction)
- Balance all atoms other than O and H
- Balance O by adding  $\text{H}_2\text{O}$
- Balance H by adding  $\text{H}^+$
- Balance charge by adding  $\text{e}^-$
- Multiply both half reactions by a factor to equate no of  $\text{e}^-$
- Add the two half-reactions and cancel out the  $\text{e}^-$ 's and any substances appearing on both side
- Double check:
  - ✓ All elements balanced
  - ✓ Charges balanced

## EXAMPLE / PRACTICE

Balance the redox reaction  $\text{Fe}^{2+}(\text{aq}) + \text{MnO}_4^{-}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq})$  in **acidic** solution, using the half-reaction method

1. **Assign oxidation states** to all atoms and identify the substances being oxidized and reduced.



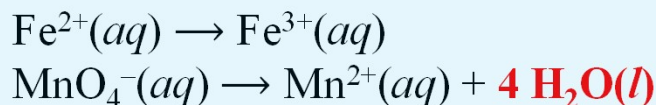
2. **Separate the overall reaction into two half-reactions**, one for oxidation and one for reduction.



3. **Balance each half-reaction with respect to mass** in the following order:

- Balance all elements other than H and O.
- Balance O by adding  $\text{H}_2\text{O}$ .

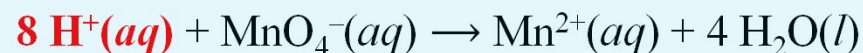
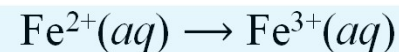
All elements other than hydrogen and oxygen are balanced, so you can proceed to the next step.



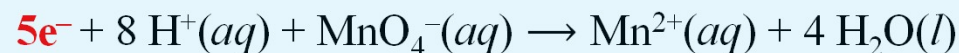
Balance the redox reaction  $\text{Fe}^{2+}(\text{aq}) + \text{MnO}_4^{-}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq})$   
in acidic solution, using the half-reaction method (*continued*)

⋮

- Balance H by adding  $\text{H}^{+}$ .



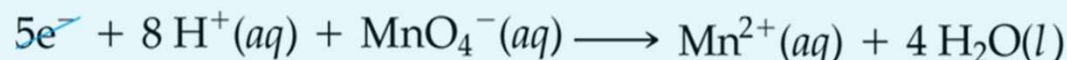
**4. Balance each half-reaction with respect to charge** by adding electrons to the right side of the oxidation half-reaction and the left side of the reduction half-reaction. (The sum of the charges on both sides of each equation should be equal.)



**5. Make the number of electrons in both half-reactions equal** by multiplying one or both half- reactions by a small whole number.



**6. Add the two half-reactions together**, canceling electrons and other species as necessary.



Balance the redox reaction  $\text{Fe}^{2+}(\text{aq}) + \text{MnO}_4^{-}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq})$   
in acidic solution, using the half-reaction method (*continued*)

⋮

**7. Verify that the reaction is balanced**  
with respect to both mass\* and charge.

| Reactants  | Products   |
|------------|------------|
| 5 Fe       | 5 Fe       |
| 8 H        | 8 H        |
| 1 Mn       | 1 Mn       |
| 4 O        | 4 O        |
| +17 charge | +17 charge |

\* Not literally “mass”. It just means the number of atoms for each element

# Balancing Redox Reactions in Basic Conditions

- Write both half reactions
- Balance all atoms other than O and H
- Balance O by adding  $\text{H}_2\text{O}$
- Balance H by adding  $\text{H}^+$
- Balance charge by adding  $\text{e}^-$
- Multiply both half reactions by a factor to equate no of  $\text{e}^-$
- Add the two half-reactions and cancel out the electrons

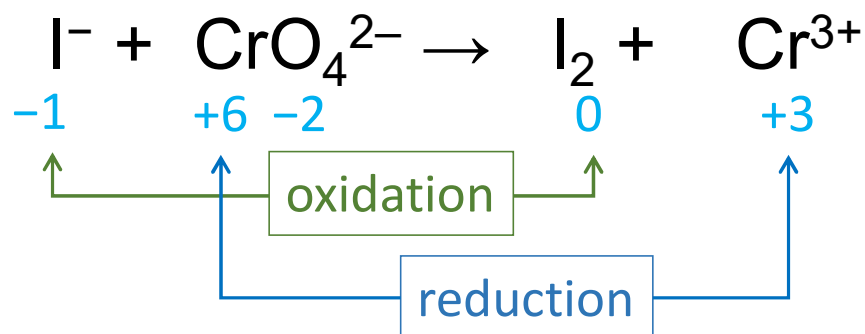
Start like acidic

- “Neutralize”  $\text{H}^+$  by adding same no. of  $\text{OH}^-$  to both sides
- Combine  $\text{H}^+$  and  $\text{OH}^-$  on the same side to make  $\text{H}_2\text{O}$
- Cancel appropriate no. of  $\text{H}_2\text{O}$  (and any other species) appearing on both sides, so that the remainder appears on only one side
- Double check: ✓ All elements balanced  
✓ Charges balanced

## Practice – Balancing redox reactions, Basic Conditions

Balance the reaction equation  $\text{I}^- + \text{CrO}_4^{2-} \rightarrow \text{I}_2 + \text{Cr}^{3+}$

assign oxidation states and  
determine element oxidized  
and element reduced



separate into oxidation &  
reduction half-reactions



Balance the reaction equation in basic solution  $\text{I}^- + \text{CrO}_4^{2-} \rightarrow \text{I}_2 + \text{Cr}^{3+}$   
(continued)

⋮

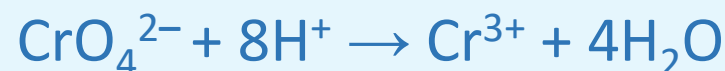
- balance atoms other than O and H



- balance O by adding  $\text{H}_2\text{O}$  to side that lacks O



- balance H by adding  $\text{H}^+$  to side that lacks H



- balance each half-reaction with respect to charge by

- adding electrons on product side for oxidation



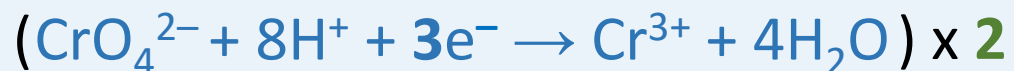
- adding electrons on reactant side for reduction



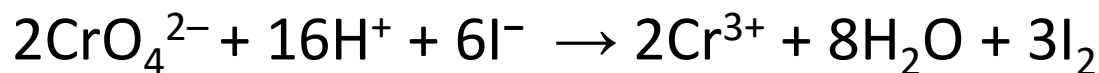
Balance the reaction equation in basic solution  $\text{I}^- + \text{CrO}_4^{2-} \rightarrow \text{I}_2 + \text{Cr}^{3+}$   
(continued)

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Multiply, if needed, each half-reaction by an appropriate factor so that both involve the same number of electrons (and they are the smallest possible integers)



Add the half reactions and cancel electrons

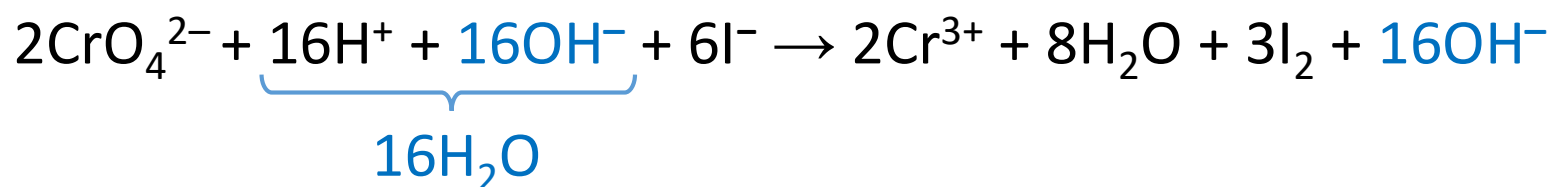


Balance the reaction equation in basic solution  $\text{I}^- + \text{CrO}_4^{2-} \rightarrow \text{I}_2 + \text{Cr}^{3+}$   
(continued)

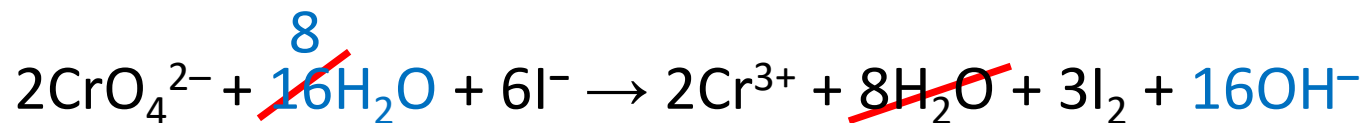
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So far it has been the same as for balancing in acidic solution

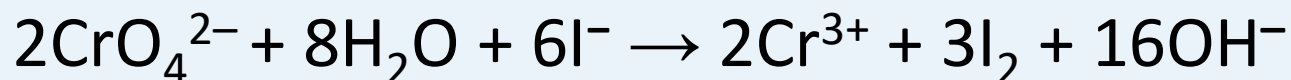
Now add enough  $\text{OH}^-$  to neutralize  $\text{H}^+$  to produce  $\text{H}_2\text{O}$ , and the same number of  $\text{OH}^-$  to the other side



Cancel  $\text{H}_2\text{O}$  (and any other species appearing on both sides at this point)



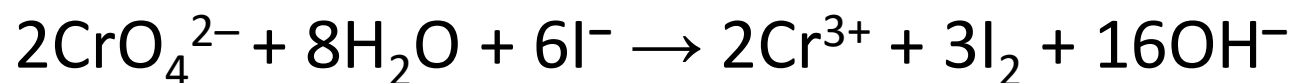
**Balanced in basic solution**



Balance the reaction equation in basic solution  $\text{I}^- + \text{CrO}_4^{2-} \rightarrow \text{I}_2 + \text{Cr}^{3+}$   
(continued)

⋮

Finally, make sure that all the elements as well as charges are balanced:



| reactant side | Element | product side |
|---------------|---------|--------------|
| 2             | Cr      | 2            |
| 6             | I       | 6            |
| 16            | O       | 16           |
| 16            | H       | 16           |
| -10           | charge  | -10          |