To assign an oxidation number (\(N_{ox}\)):
- The oxidation number of an elemental substance is zero. Examples: Na, Cl, Ne, Fe
- The oxidation number for a monatomic ion is equal to the charge on that ion. Examples: \(Na^+, Cl^-, Fe^{2+}, Fe^{3+}\)
- The sum of the oxidation numbers of all atoms in a species must equal the total charge on that species. Examples: for \(CO_2, N_{ox}C + 2 \cdot N_{ox}O = 0\); for \(CO_2^{-}, N_{ox}C + 3 \cdot N_{ox}O = -2\)
- Fluorine in polyatomic species is assigned \(N_{ox} = -1\).
- Hydrogen in polyatomic species is assigned \(N_{ox} = +1\) with nonmetals (covalent or molecular compounds) and \(N_{ox} = -1\) with metals (ionic compounds - hydride ion is \(H^-\)).
- Oxygen in polyatomic species is usually assigned \(N_{ox} = -2\). Exceptions include compounds of O with F (assign F an \(N_{ox}\) of \(-1\) first, then determine \(N_{ox}\) of O); peroxides where oxygen is present as the peroxide anion, \(O_2^{2-}\), with \(N_{ox} = -1\) (example: \(H_2O_2\)); superoxides where oxygen is present as the superoxide anion, \(O_2^-\), with \(N_{ox} = -1/2\) (example: \(KO_2\)).

Steps for balancing equations for reactions taking place in acidic medium:

Consider the reaction between permanganate ion and oxalic acid to form manganese (II) ion and carbon dioxide. Write the overall balanced redox equation for this reaction.

1. Write the skeletal equation for the reaction; identify what is oxidized and reduced.
   \[\text{MnO}_4^- + \text{H}_2\text{C}_2\text{O}_4 \rightarrow \text{Mn}^{2+} + \text{CO}_2\]
   C is oxidized, +3 \(\rightarrow\) +4 \(\cdot\) \(\text{H}_2\text{C}_2\text{O}_4\) is the reducing agent
   Mn is reduced, +7 \(\rightarrow\) +2 \(\cdot\) \(\text{MnO}_4^-\) is the oxidizing agent

2. Write the skeletal half reactions for the oxidation and reduction.
   oxidation: \(\text{H}_2\text{C}_2\text{O}_4 \rightarrow \text{CO}_2\)
   reduction: \(\text{MnO}_4^- \rightarrow \text{Mn}^{2+}\)

3. Balance elements other than H and O.
   oxidation: \(\text{H}_2\text{C}_2\text{O}_4 \rightarrow 2 \text{CO}_2\)
   reduction: \(\text{MnO}_4^- \rightarrow \text{Mn}^{2+}\)

4. Balanced O by adding \(\text{H}_2\text{O}\).
   oxidation: \(\text{H}_2\text{C}_2\text{O}_4 \rightarrow 2 \text{CO}_2\)
   reduction: \(\text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}\)

5. Balance H by adding \(\text{H}^+\).
   oxidation: \(\text{H}_2\text{C}_2\text{O}_4 \rightarrow 2 \text{CO}_2 + 2 \text{H}^+\)
   reduction: \(8 \text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}\)

6. Balance charge by adding electrons (e\(^-\)).
   oxidation: \(\text{H}_2\text{C}_2\text{O}_4 \rightarrow 2 \text{CO}_2 + 2 \text{H}^+ + 2 \text{e}^-\)
   reduction: \(5 \text{e}^- + 8 \text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}\)

*At this point the individual half reactions are balanced.*

7. Prepare to add the half reactions together by multiplying by an appropriate factor so that the electrons cancel out.
   \[5 \times \{ \text{oxidation: } \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2 \text{CO}_2 + 2 \text{H}^+ + 2 \text{e}^-\} \]
   \[2 \times \{ \text{reduction: } 5 \text{e}^- + 8 \text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O} \} \]
   oxidation: \(5 \text{H}_2\text{C}_2\text{O}_4 \rightarrow 10 \text{CO}_2 + 10 \text{H}^+ + 10 \text{e}^-\)
   reduction: \(10 \text{e}^- + 16 \text{H}^+ + 2 \text{MnO}_4^- \rightarrow 2 \text{Mn}^{2+} + 8 \text{H}_2\text{O}\)

8. Add 1/2 reactions together, simplify (cancel species that are the same on both sides of the equation) ; add physical states.
   \[5 \text{H}_2\text{C}_2\text{O}_4 + 10 \text{e}^- + 16 \text{H}^+ + 2 \text{MnO}_4^- \rightarrow 10 \text{CO}_2 + 10 \text{H}^+ + 10 \text{e}^- + 2 \text{Mn}^{2+} + 8 \text{H}_2\text{O}\]
   cancel \(\text{e}^-\), reduce \(\text{H}^+\)
   \[5 \text{H}_2\text{C}_2\text{O}_4 \text{(aq)} + 6 \text{H}^+ \text{(aq)} + 2 \text{MnO}_4^- \text{(aq)} \rightarrow 10 \text{CO}_2 \text{(g)} + 2 \text{Mn}^{2+} \text{(aq)} + 8 \text{H}_2\text{O} \text{(l)}\]
Steps for balancing equations for reactions taking place in basic medium:
Consider the reaction between dichlorine heptaoxide gas and hydrogen peroxide to form chlorite ion and oxygen gas.

1. Write the skeletal equation for the reaction; identify what is oxidized and reduced.
   \[ \text{Cl}_2\text{O}_7 + \text{H}_2\text{O}_2 \rightarrow \text{ClO}_2^- + \text{O}_2 \]
   O is oxidized, \(-1 \rightarrow 0\) :: \(\text{H}_2\text{O}_2\) is the reducing agent
   Cl is reduced, \(+7 \rightarrow +3\) :: \(\text{Cl}_2\text{O}_7\) is the oxidizing agent

2. Write the skeletal half reactions for the oxidation and reduction.
   oxidation: \(\text{H}_2\text{O}_2 \rightarrow \text{O}_2\)
   reduction: \(\text{Cl}_2\text{O}_7 \rightarrow \text{ClO}_2^-\)

3. Balance elements other than H and O.
   oxidation: \(\text{H}_2\text{O}_2 \rightarrow \text{O}_2\)
   reduction: \(\text{Cl}_2\text{O}_7 \rightarrow 2 \text{ClO}_2^-\)

4. Balanced O by adding \(\text{H}_2\text{O}\).
   oxidation: \(\text{H}_2\text{O}_2 \rightarrow \text{O}_2\)
   reduction: \(\text{Cl}_2\text{O}_7 \rightarrow 2 \text{ClO}_2^- + 3 \text{H}_2\text{O}\)

5. Balance H by adding \(\text{H}^+\). Then, add that same number of \(\text{OH}^-\) on both sides of the half reaction. Simplify by recognizing that \(\text{H}^+ + \text{OH}^- = \text{H}_2\text{O}\)
   oxidation: \(2 \text{OH}^- + \text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2 \text{H}^+ + 2 \text{OH}^-\)
   reduction: \(6 \text{OH}^- + 6 \text{H}^+ + \text{Cl}_2\text{O}_7 \rightarrow 2 \text{ClO}_2^- + 3 \text{H}_2\text{O} + 6 \text{OH}^-\)
   oxidation: \(2 \text{OH}^- + \text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2 \text{H}_2\text{O}\)
   reduction: \(6 \text{H}_2\text{O} + \text{Cl}_2\text{O}_7 \rightarrow 2 \text{ClO}_2^- + 3 \text{H}_2\text{O} + 6 \text{OH}^-\)

6. Balance charge by adding electrons \(\text{e}^-\).
   oxidation: \(2 \text{OH}^- + \text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2 \text{H}_2\text{O} + 2 \text{e}^-\)
   reduction: \(8 \text{e}^- + 6 \text{H}_2\text{O} + \text{Cl}_2\text{O}_7 \rightarrow 2 \text{ClO}_2^- + 3 \text{H}_2\text{O} + 6 \text{OH}^-\)

* At this point the individual half reactions are balanced *

7. Prepare to add the half reactions together by multiplying by an appropriate factor so that the electrons cancel out.
   \[4 \times \text{oxidation: } 2 \text{OH}^- + \text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2 \text{H}_2\text{O} + 2 \text{e}^- \]
   \[1 \times \text{reduction: } 8 \text{e}^- + 6 \text{H}_2\text{O} + \text{Cl}_2\text{O}_7 \rightarrow 2 \text{ClO}_2^- + 3 \text{H}_2\text{O} + 6 \text{OH}^-\]
   oxidation: \(8 \text{OH}^- + 4 \text{H}_2\text{O}_2 \rightarrow 4 \text{O}_2 + 8 \text{H}_2\text{O} + 8 \text{e}^-\)
   reduction: \(8 \text{e}^- + 6 \text{H}_2\text{O} + \text{Cl}_2\text{O}_7 \rightarrow 2 \text{ClO}_2^- + 3 \text{H}_2\text{O} + 6 \text{OH}^-\)

8. Add 1/2 reactions together, simplify (cancel species that are the same on both sides of the equation) ; add physical states.
   \(8 \text{OH}^- + 4 \text{H}_2\text{O}_2 + 8 \text{e}^- + 6 \text{H}_2\text{O} + \text{Cl}_2\text{O}_7 \rightarrow 4 \text{O}_2 + 8 \text{H}_2\text{O} + 8 \text{e}^- + 2 \text{ClO}_2^- + 3 \text{H}_2\text{O} + 6 \text{OH}^-\)
   cancel \(\text{e}^-\), reduce \(\text{H}_2\text{O}\) and \(\text{OH}^-\)
   \(2 \text{OH}^- (\text{aq}) + 4 \text{H}_2\text{O}_2 (\text{aq}) + \text{Cl}_2\text{O}_7 (\text{g}) \rightarrow 4 \text{O}_2 (\text{g}) + 2 \text{ClO}_2^- (\text{aq}) + 5 \text{H}_2\text{O} (\text{l})\)