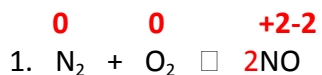


Practice Balancing Redox Reactions

KEY



1. Assign Oxidation #'s

2. Determine which element is oxidized and which is reduced...

Oxidized: **N**

Reduced: **O**

3. Write out the half-reactions. (*Hint: Are your half-reactions balanced?*)

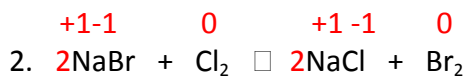
Oxidation: $\text{N}_2 \rightarrow 2\text{N}^{+2} + 4\text{e}^-$

Reduction: $\text{O}_2 + 4\text{e}^- \rightarrow 2\text{O}^{2-}$

4. Balance the electrons...**Done!**

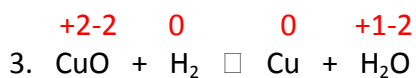
5. Transfer the new balancing coefficients back into the original...**The 2's were added above.**

6. Balance the remaining elements...**Didn't need to...**



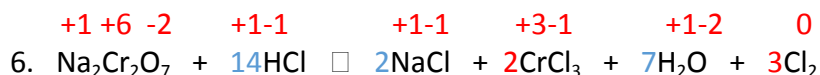
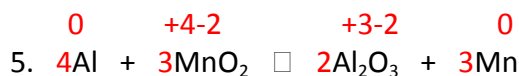
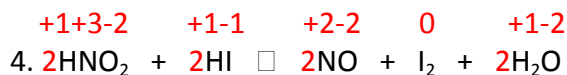
Ox: $2\text{Br}^{-1} \rightarrow \text{Br}_2 + 2\text{e}^-$

Red: $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^{-1}$

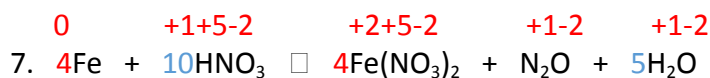


Ox: $\text{H}_2 \rightarrow 2\text{H}^{+1} + 2\text{e}^-$

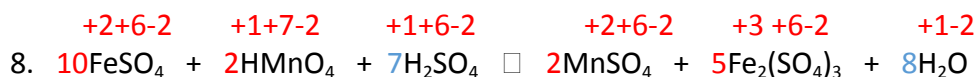
Red: $\text{Cu}^{+2} + 2\text{e}^- \rightarrow \text{Cu}$



*Not all of the Cl⁻ were oxidized, so the balancing on the HCl had to be adjusted to account for them. Blue are coefficients added to balance non-redox elements.



*Not all of the N⁺⁵ were oxidized, so the balancing on the HNO₃ had to be adjusted to account for them. Blue are coefficients added to balance non-redox elements.



Blue are coefficients added to balance non-redox elements.

