

Balancing Redox Reactions

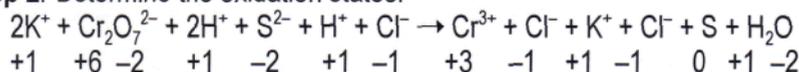
Redox equations are often too complex to balance by inspection alone. Instead, they are balanced by the *half-reaction method* or *ion-electron method*. In redox reactions, the number of electrons lost is always equal to the number of electrons gained. Keeping track of the electrons helps to balance the parts of the equation that can't be balanced by inspection. This is done by the procedure outlined below.

Balance the following: $\text{K}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{S} + \text{HCl} \rightarrow \text{CrCl}_3 + \text{KCl} + \text{S} + \text{H}_2\text{O}$

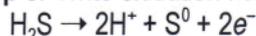
Step 1: Write the ionic equation.



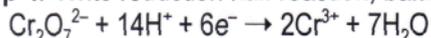
Step 2: Determine the oxidation states.



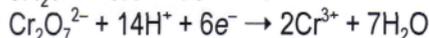
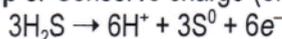
Step 3: Write oxidation half reaction, balancing atoms and charge.



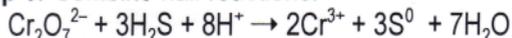
Step 4: Write reduction half reaction, balancing atoms and charge.



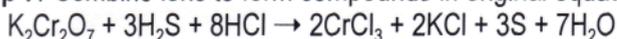
Step 5: Conserve charge (electrons lost = electrons gained).



Step 6: Combine half reactions.



Step 7: Combine ions to form compounds in original equation.



REDOX Inc.

The record keeping around here is horrific!



In **Step 1** the ions are separated making the spectators easier to identify. In **Step 2** the oxidation states are determined so it is possible to tell what was oxidized and what was reduced. In **Steps 3 and 4**, half reactions are written showing the number of electrons transferred. Note that in the oxidation half, 2H^+ are needed to balance the hydrogen in hydrogen sulfide. In the reduction half, $7\text{H}_2\text{O}$ are needed to balance the oxygen

in the dichromate ion, and as a result 14H^+ are needed on the reactant side. In **Step 5**, the half reactions are multiplied by the correct coefficients to make the number of electrons lost equal the number of electrons gained. In **Step 6**, note that the H^+ ions remaining are the net from the two half reactions where they are on opposite sides of the equation.

Balance the equations below by following the procedure above.

- $\text{H}_2\text{S}(aq) + \text{HNO}_3(aq) \rightarrow \text{S}(s) + \text{NO}_2(g) + \text{H}_2\text{O}(l)$
- $\text{LiNO}_3(aq) + \text{FeCl}_2(aq) + \text{HCl}(aq) \rightarrow \text{NO}(g) + \text{LiCl}(aq) + \text{FeCl}_3(aq) + \text{H}_2\text{O}(l)$
- $\text{Na}_2\text{Cr}_2\text{O}_7(aq) + \text{HI}(aq) \rightarrow \text{CrI}_3(aq) + \text{NaI}(aq) + \text{I}_2(s) + \text{H}_2\text{O}(l)$
- $\text{NaClO}_3 + \text{HCl} \rightarrow \text{ClO}_2 + \text{NaClO}_4 + \text{NaCl} + \text{H}_2\text{O}$
- $\text{PbS}(s) + \text{HNO}_3(aq) \rightarrow \text{Pb}(\text{NO}_3)_2(aq) + \text{S}(s) + \text{NO}(g) + \text{H}_2\text{O}(l)$

Determining the Voltage of Electrochemical Cells

Chemical reactions often involve the movement of electrons. The driving force that moves the electrons can be measured. It is the voltage. The voltage of an electrochemical cell can be determined using the *Standard Reduction Table*.

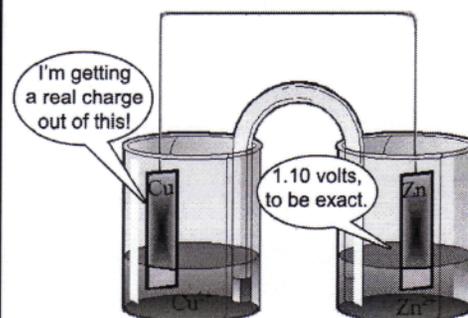
Procedure

1. All half reactions on the *Standard Reduction Potentials Table* are compared to hydrogen ($E^{\circ} = 0$)
2. All half reactions can be read in reverse as oxidations in which case the sign of the voltage, E° , is changed
3. The net voltage is the sum of the voltages of the oxidation half reactions and the reduction half reactions (see chart)

Example

What voltage is associated with the reaction $\text{CuSO}_4 + \text{Zn} \rightarrow \text{ZnSO}_4 + \text{Cu}$?

		$\text{Zn}^0 \rightarrow \text{Zn}^{2+} + 2e^-$	$E^{\circ} = 0.76\text{v}$
$\text{Cu}^{2+} + 2e^-$	\rightarrow	Cu^0	$E^{\circ} = 0.34\text{v}$
$\text{Cu}^{2+} + \text{Zn}^0$	\rightarrow	$\text{Zn}^{2+} + \text{Cu}^0$	$E^{\circ} = 1.10\text{v}$



Existential discussions in voltaic cells

Write the half reactions for each of the following reactions, balance them, and determine the voltage (E°) associated with the reaction by using the *Standard Reduction Table*.

1. $\text{Cu} + \text{AgNO}_3 \rightarrow \text{Ag} + \text{Cu}(\text{NO}_3)_2$
2. $\text{K}_2\text{Cr}_2\text{O}_7 + \text{SnCl}_2 + \text{HCl} \rightarrow \text{CrCl}_3 + \text{SnCl}_4 + \text{KCl} + \text{H}_2\text{O}$
3. $\text{SnCl}_2 + \text{HgCl}_2 \rightarrow \text{SnCl}_4 + \text{Hg}_2\text{Cl}_2$
4. $\text{Sn} + \text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SnO}_3 + \text{NO}$
5. $\text{KBr} + \text{Fe}_2(\text{SO}_4)_3 \rightarrow \text{Br}_2 + \text{K}_2\text{SO}_4 + \text{FeSO}_4$
6. $\text{Fe} + \text{CuSO}_4 \rightarrow \text{Cu} + \text{Fe}_2(\text{SO}_4)_3$
7. $\text{KMnO}_4 + \text{HCl} \rightarrow \text{KCl} + \text{MnCl}_2 + \text{H}_2\text{O} + \text{Cl}_2$
8. $\text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2$
9. $\text{HBr} + \text{MnO}_2 \rightarrow \text{MnBr}_2 + \text{H}_2\text{O} + \text{Br}_2$
10. $\text{HCl} + \text{K}_2\text{SO}_4 \rightarrow \text{KCl} + \text{SO}_2 + \text{H}_2\text{O} + \text{Cl}_2$