("q&d") Redox Reactions Guide for CH 223

"q&d" = "Quick 'n' Dirty"

<u>Part I</u>: General Instructions - *for reactions where pH is not important*

1) Recognize that the reaction is an oxidation-reduction reaction. Look for changes in oxidation number on similar atoms in both the reactant and product side.

Example: $HCl + NaOH \rightarrow H_2O_{(l)} + NaCl$ *not an oxidation-reduction reaction*
Example: $Ag^+ + Cu_{(s)} \rightarrow Ag_{(s)} + Cu^{2+}$ *oxidation-reduction*

2) Separate the process into two **half-reactions**. One side will be a **reduction** and the other will be an **oxidation**.

Example: $Ag^+ \rightarrow Ag_{(s)}$
$Cu_{(s)} \rightarrow Cu^{2+}$

3) Balance each half-reaction for **mass** (atoms) and **charge** (using electrons).

Example: $Ag^+ + e^- \rightarrow Ag_{(s)}$ *reduction*
$Cu_{(s)} \rightarrow Cu^{2+} + 2\ e^-$ *oxidation*

4) Multiply each half-reaction by a factor that makes the number of electrons equal.

Example: $Ag^+ + e^- \rightarrow Ag_{(s)}$
$Cu_{(s)} \rightarrow Cu^{2+} + 2\ e^-$
If we multiply the first reaction by **2** and the second reaction by **1**, the number of electrons in each half-reaction will be equal.
$2\ Ag^+ + 2\ e^- \rightarrow 2\ Ag_{(s)}$
$Cu_{(s)} \rightarrow Cu^{2+} + 2\ e^-$

5) Add the two half-reactions to create the overall balanced equation. The equation should be balanced for mass and devoid of electrons.

Example: $2\ Ag^+ + \cancel{2\ e^-} \rightarrow 2\ Ag_{(s)}$
$Cu_{(s)} \rightarrow Cu^{2+} + \cancel{2\ e^-}$

$\mathbf{2\ Ag^+ + Cu_{(s)} \rightarrow 2\ Ag_{(s)} + Cu^{2+}}$

6) Confirm that mass and charge are balanced in the overall equation.

Example: 2 silver atoms on reactant and product side
1 copper atom on reactant and product side
mass balanced
2(+1) = +2 charge on reactant side
+2 charge on product side
charge balanced

<u>Part II</u>: Acidic Conditions - *for reactions where pH < 7*

1) Recognize that the reaction is an oxidation-reduction reaction. Look for changes in oxidation number on similar atoms in both the reactant and product side.

Example: $VO_2^+ + Zn_{(s)} \rightarrow VO^{2+} + Zn^{2+}$ *oxidation-reduction*

2) Separate the process into two **half-reactions**. One side will be a **reduction** and the other will be an **oxidation**.

Example: $VO_2^+ \rightarrow VO^{2+}$
$Zn_{(s)} \rightarrow Zn^{2+}$

3) Balance each half-reaction for **mass** (atoms) and **charge** (using electrons). Note that in *acidic solution*, mass can be balanced by using water and H^+ if appropriate. Water will go to the side that is *oxygen deficient*, and hydrogen ions will go to the side that is *hydrogen deficient*. Some half-reactions will not need any water or hydrogen ions.

Example: $2\ H^+ + VO_2^+ + e^- \rightarrow VO^{2+} + H_2O_{(l)}$ *reduction*
$Zn_{(s)} \rightarrow Zn^{2+} + 2\ e^-$ *oxidation*

4) Multiply each half-reaction by a factor that makes the number of electrons equal.

Example: $2\ H^+ + VO_2^+ + e^- \rightarrow VO^{2+} + H_2O_{(l)}$
$Zn_{(s)} \rightarrow Zn^{2+} + 2\ e^-$
If we multiply the first reaction by **2** and the second reaction by **1**, the number of electrons in each half-reaction will be equal.
$4\ H^+ + 2\ VO_2^+ + 2\ e^- \rightarrow 2\ VO^{2+} + 2\ H_2O_{(l)}$
$Zn_{(s)} \rightarrow Zn^{2+} + 2\ e^-$

5) Add the two half-reactions to create the overall balanced equation. The equation should be balanced for mass & devoid of electrons. Cancel water and/or protons if present on both sides of the equation.

Example: $4\ H^+ + 2\ VO_2^+ + \cancel{2\ e^-} \rightarrow 2\ VO^{2+} + 2\ H_2O_{(l)}$
$Zn_{(s)} \rightarrow Zn^{2+} + \cancel{2\ e^-}$

$\mathbf{4\ H^+ + 2\ VO_2^+ + Zn_{(s)} \rightarrow 2\ VO^{2+} + 2\ H_2O_{(l)} + Zn^{2+}}$

6) Confirm that mass and charge are balanced in the overall equation.

Example: 4 hydrogen atoms on each side
2 vanadium atoms on each side
4 oxygen atoms on each side
1 zinc atom on each side
mass balanced
4(+1) + 2(+1) = +6 charge on reactant side
2(+2) +2 = +6 charge on product side
charge balanced

Part III: Basic Conditions - *for reactions where pH > 7*

1) Recognize that the reaction is an oxidation-reduction reaction. Look for changes in oxidation number on similar atoms in both the reactant and product side.

Example: $MnO_4^- + HO_2^- \rightarrow MnO_4^{2-} + O_2$ *oxidation-reduction*

2) Separate the process into two **half-reactions**. One will be a **reduction** and the other an **oxidation**.

Example: $MnO_4^- \rightarrow MnO_4^{2-}$
$HO_2^- \rightarrow O_2$

3) Balance each half-reaction for **mass** (atoms) and **charge** (using electrons). Note that in *basic solution*, mass can be balanced by using water and OH^- if appropriate. Hydroxide will go to the side that is *oxygen deficient*, and water will go to the side that is *hydrogen deficient*. Some half-reactions will not need any water or hydrogen ions.

Example: $MnO_4^- + e^- \rightarrow MnO_4^{2-}$ *reduction*
$OH^- + HO_2^- \rightarrow O_2 + 2\,e^- + H_2O_{(l)}$ *oxidation*

4) Multiply each half-reaction by a factor that makes the number of electrons equal.

Example: $MnO_4^- + e^- \rightarrow MnO_4^{2-}$
$OH^- + HO_2^- \rightarrow O_2 + 2\,e^- + H_2O_{(l)}$
If we multiply the first reaction by **2** and the second reaction by **1**, the number of electrons in each half-reaction will be equal.
$2\,MnO_4^- + 2\,e^- \rightarrow 2\,MnO_4^{2-}$
$OH^- + HO_2^- \rightarrow O_2 + 2\,e^- + H_2O_{(l)}$

5) Add the two half-reactions to create the overall balanced equation. The equation should be balanced for mass and devoid of electrons. Cancel water and/or hydroxide if present on both sides of the equation.

Example: $2\,MnO_4^- + \cancel{2\,e^-} \rightarrow 2\,MnO_4^{2-}$
$OH^- + HO_2^- \rightarrow O_2 + \cancel{2\,e^-} + H_2O_{(l)}$

$\mathbf{2\,MnO_4^- + OH^- + HO_2^- \rightarrow 2\,MnO_4^{2-} + O_2 + H_2O_{(l)}}$

6) Confirm that mass and charge are balanced in the overall equation.

Example: 2 manganese atoms on each side
11 oxygen atoms on each side
2 hydrogen atoms on each side - *mass balanced*
2(-1) + (-1) + (-1) = -4 charge on reactant side
2(-2) = -4 charge on product side - *charge balanced*

7) *Alternatively*, you may balance basic redox reactions using the acidic process used in Part II. Upon completing step 5, add OH^- to neutralize any H^+ present. Add a similar amount of OH^- to both sides. All of the H^+ will be converted to H_2O, and the opposite side should have OH^- present.