Lesson 2 – Redox. Writing oxidation and reduction half equations for more complex oxidants and reductants.

Simple oxidants such as Ag^+ are somewhat easy to write a balanced half equation for.

$Ag^+(aq) + e \rightarrow Ag(s)$

$Sn^{4+}(aq) + 4e \rightarrow Sn(s)$

$Pb(s) \rightarrow Pb^{2+}(aq) + 2e$

Electrons always added to the most positive side

$Cr_2O_7^{-2} / Cr^{3+}$

To write a balanced half equation follow the steps

1. Balance for all elements other than O or H

$Cr_2O_7^{-2} \rightarrow 2Cr^{3+}$

2. balance for oxygen by adding water to the side with least number of oxygens

$Cr_2O_7^{-2} \rightarrow 2Cr^{3+} + 7H_2O$

3. balance for hydogens by adding $H^+(aq)$ to the side with least hydrogens

$14H^+(aq) + Cr_2O_7^{-2} \rightarrow 2Cr^{3+} + 7H_2O$

4. Balance for charge by adding electrons to the most positive side

$6e + 14H^+(aq) + Cr_2O_7^{-2} \rightarrow 2Cr^{3+} + 7H_2O$

1. Use the steps outlined above to give a balanced half equation, states not included, for the reaction of the the following conjugate pairs. Indicate if the reaction is an oxidation or a reduction.
Keep in mind : Reduction – electrons appear on the left
Oxidation – electrons appear on the right

a. MnO_4^- / Mn^{2+}

b. BiO_4^{-3} / Bi^{3+}

c. SO_3^{-2} / SO_4^{-2}

d. NO_3^- / NO_2

e. $C_2O_4^{-2} / CO_2$

f. H_5IO_6 / IO_3^-

g. BrO_3^{-3} / Br^-

h. MnO_2 / Mn_2O_3

i. $HClO_4 / HCl$

j. P_4H_{10} / P_2O_5

k. P_4 / PO_4^{-3}