Balancing Redox Equations Practice

Remember the rules:
a) determine the oxidation numbers. Identify the reduction and the oxidation half rxns.
b) write the two half reactions separately in net ionic equation form (place $e^-$'s correctly)
c) balance each half reaction separately. If acidic conditions, balance O with $H_2O$ and H with $H^+$. If basic, balance equations as in the acidic case. Then add $OH^-$'s to both sides equal to the number of $H^+$'s present to neutralize those to $H_2O$. Simplify afterwards.
d) Add the two half reactions, making sure to multiply each by some factor so e's cancel out in the end.

ACID SOLUTION:

1. $I_2 + H_2S$ - - - > $H^+ + I^- + S(s)$
2. $I^- + H_2SO_4$ - - - > $I_2 + SO_2$
3. $Ag + NO_3^-$ - - - > $Ag^+ + NO$
4. $CuS + NO_3^-$ - - - > $Cu^{2+} + SO_4^{2-} + NO$
5. $S_2O_3^{2-} + I_2$ - - - > $I^- + S_4O_6^{2-}$
6. $Zn + NO_3^-$ - - - > $Zn^{2+} + NH_4^+$
7. $HS_2O_3^-$ - - - > $S + HSO_4^-$
8. $ClO_3^- + As_2S_3$ - - - > $Cl^- + H_2AsO_4^- + SO_4^{2-}$
9. $MnO_4^{2-}$- - - > $MnO_2 + MnO_4^-$
10. $I_2 + IO_4^- + Cl^{-=}$ - - - > $ICl_2^-$

BASIC SOLUTION:

1. $AL + NO_3^- + OH^-$ - - - > $Al(OH)_4^- + NH_3$
2. $PbO_2 + Cl^-$ - - - > $ClO^- + Pb(OH)_3^-$
3. $N_2H_4 + Cu(OH)_2$ - - - > $N_2 + Cu$
4. $Ag_2S + CN^- + O_2$ - - - > $S + Ag(CN)_2^-$
5. $ClO^- + Fe(OH)_3$ - - - > $Cl^- + FeO_4^{2-}$
6. $HO_2^- + Cr(OH)_3^-$ - - - > $CrO_4^{2-} + OH^-$
7. $Cu(NH_3)_4^{2+} + S_2O_4^{2-}$ - - - > $SO_3^{2-} + Cu + NH_3$
8. $ClO_2 + OH^-$ - - - > $ClO_2^- + ClO_3^-$
9. $V + H_2O$ - - - > $HV_6O_{17}^{3-} + H_2$
10. $I_2 + IO_4^- + Cl^{-=}$ - - - > $ICl_2$