Lesson 1 – oxidation, reduction and redox reactions.

1. For each of the following reactions indicate the following:
   - type of reaction
   - identify the reductant or oxidant
   - identify the conjugate.

   The fist one is done for you.

   a. $Fe^{2+}(aq) + e \rightarrow Fe^{+}(aq)$
      - type of reaction = reduction (electrons are used to reduce the charge on the oxidant)
      - Oxidant is $Fe^{2+}(aq)$ (it accepts electrons)
      - Conjugate reductant = $Fe^{+}(aq)$

   b. $2H^{+}(aq) + 2e \rightarrow H_2(g)$
      - type of reaction = *reduction (electrons are being used to reduce the oxidation number of H^{+} to zero)*
      - identify the reductant or oxidant = *$H^{+}(aq)$ is the oxidant (it accepts electrons)*
      - identify the conjugate = *$H_2(g)$ is the conjugate reductant*

   c. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e$
      - type of reaction = *oxidation (electrons are being lost to increase the oxidation number of Zn from zero to 2+)*
      - identify the reductant or oxidant = *Zn(s) is a reductant (it gives away electrons)*
      - identify the conjugate = *$Zn^{2+}(aq)$ conjugate oxidant*

   d. $2Br^{-}(aq) \rightarrow 2e + Br_2(l)$
      - type of reaction = *oxidation (electrons are being lost to increase the oxidation number of Br^{-} from -1 to 0)*
      - identify the reductant or oxidant = *$Br^{-}aq)$ is a reductant (it gives away electrons)*
      - identify the conjugate = *$Br_2(l)$ conjugate oxidant*

   e. $Sn^{4+}(aq) + 2e \rightarrow Sn^{2+}(aq)$
      - type of reaction = *reduction (electrons are being used to decrease the oxidation number of Sn^{4+} from 4+ to 2+)*
      - identify the reductant or oxidant = *$Sn^{4+}(aq)$ is an oxidant (it accepts electrons)*
      - identify the conjugate = *$Sn^{2+}(aq)$ conjugate reductant*

   f. $K^{+}(aq) + e \rightarrow K(s)$
      - type of reaction = *reduction (electrons are being used to decrease the oxidation number of K^{+} from 1+ to 0)*
      - identify the reductant or oxidant = *$K^{+}(aq)$ is an oxidant (it accepts electrons)*
      - identify the conjugate = *K(s) is the conjugate reductant*

   g. $Mn(s) \rightarrow Mn^{2+}(aq) + 2e \rightarrow$
      - type of reaction = *Oxidation (electrons are being lost to increase the oxidation number of Mn(s) from 0 to 2+)*
      - identify the reductant or oxidant = *Mn(s) is an reductant (it donates electrons)*
      - identify the conjugate = *$Mn^{2+}(aq)$ conjugate oxidant*

2. Consider the following balanced equations representing different redox reactions.
Write the balanced ionic equations for each.
The first one is done for you.

$AgNO_3(aq) + Fe(s) \rightarrow Ag(s) + Fe(NO_3)_2(aq)$

$Ag^+(aq) + NO_3^-(aq) + Fe(s) \rightarrow Ag(s) + Fe^{2+}(aq) + 2NO_3^-(aq)$

$2Ag^+(aq) + 2NO_3^-(aq) + Fe(s) \rightarrow 2Ag(s) + Fe^{2+}(aq) + 2NO_3^-(aq)$

$2Ag^+(aq) + Fe(s) \rightarrow 2Ag(s) + Fe^{2+}(aq)$

Make sure all ionic equations are balanced for charge and elements.

a. $Sn(NO_3)_2(aq) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Sn(NO_3)_4(aq)$

$Sn^{2+}(aq) + 2NO_3^-(aq) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow 2Ag(s) + Sn^{4+}(aq) + 4NO_3^-(aq)$

$\Rightarrow Sn^{2+}(aq) + 2Ag^+(aq) \rightarrow 2Ag(s) + Sn^{4+}(aq)$

b. $Ca(s) + Zn(NO_3)_2(aq) \rightarrow Zn(s) + Ca(NO_3)_2(aq)$

$Ca(s) + Zn^{2+}(aq) + 2NO_3^-(aq) \rightarrow Zn(S) + Ca^{2+}(aq) + 2NO_3^-(aq)$

$\Rightarrow Ca(s) + Zn^{2+}(aq) \rightarrow Zn(s) + Ca^{2+}(aq)$

c. $Al(s) + 3AgNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3Ag(s)$

$Al(s) + 3Ag^+(aq) + 3NO_3^-(aq) \rightarrow Al^{3+}(aq) + 3NO_3^-(aq) + 3Ag(s)$

$\Rightarrow Al(s) + 3Ag^+(aq) \rightarrow Al^{3+}(aq) + 3Ag(s)$

d. $2Fe(NO_3)_3(aq) + Sn(NO_3)_2(aq) \rightarrow Sn(NO_3)_4(aq) + 2Fe(NO_3)_2(aq)$

$2Fe^{3+}(aq) + 6NO_3^-(aq) + Sn^{2+}(aq) + 2NO_3^-(aq) \rightarrow Sn^{4+}(aq) + 4NO_3^-(aq) + 2Fe^{2+}(aq) + 4NO_3^-(aq)$

$\Rightarrow 2Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2Fe^{2+}(aq)$

e. $Cu(NO_3)_2(aq) + Fe(s) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$

$Cu^{2+}(aq) + 2NO_3^-(aq) + Fe(s) \rightarrow Fe^{2+}(aq) + 2NO_3^-(aq) + Cu(s)$

$\Rightarrow Cu^{2+}(aq) + Fe(s) \rightarrow Fe^{2+}(aq) + Cu(s)$

f. $Pb(s) + Sn(NO_3)_4(aq) \rightarrow Pb(NO_3)_2(aq) + Sn(NO_3)_2(aq)$

$\Rightarrow Pb(s) + Sn^{4+}(aq) + 4NO_3^-(aq) \rightarrow Pb^{2+}(aq) + 2NO_3^-(aq) + Sn^{2+}(aq) + 2NO_3^-(aq)$

$\Rightarrow Pb(s) + Sn^{4+}(aq) \rightarrow Pb^{2+}(aq) + Sn^{2+}(aq)$

g. $Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

$\Rightarrow Mg(s) + Cu^{2+}(aq) + SO_4^{-2}(aq) \rightarrow Mg^{2+}(aq) + SO_4^{-2}(aq) + Cu(s)$

$\Rightarrow Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

3. For each ionic equation mentioned in question 2. Identify the :
   i. reductant and its conjugate
   ii. Oxidant and its conjugate
   The first one is done for you
   Don't forget look for the:
   - species that accepts electrons to decrease in oxidation number (charge number)- that is the oxidant
   - species that donates electrons to increase in oxidation number - that is the reductant

   $2Ag^{+}(aq) + Fe(s) \rightarrow 2Ag(s) + Fe^{2+}(aq)$
   Ag^{+}(aq) (oxidant) / Ag(s) (conjugate reductant)
   Fe(s) (reductant) / Fe^{2+}(aq) (conjugate oxidant)

   a. $Sn^{2+}(aq) + 2Ag^{+}(aq) \rightarrow 2Ag(s) + Sn^{4+}(aq)$
   Ag^{+}(aq) (oxidant) / Ag(s) (conjugate reductant)
   Sn^{2+}(aq) (reductant) / Sn^{4+}(aq) (conjugate oxidant)

   b. $Ca(s) + Zn^{2+}(aq) \rightarrow Zn(s) + Ca^{2+}(aq)$
   Zn^{+2}(aq) (oxidant) / Zn(s) (conjugate reductant)
   Ca(s) (reductant) / Ca^{2+}(aq) (conjugate oxidant)

   c. $Al(s) + 3Ag^{+}(aq) \rightarrow Al^{3+}(aq) + 3Ag(s)$
   Ag^{+}(aq) (oxidant) / Ag(s) (conjugate reductant)
   Al(s) (reductant) / Al^{3+}(aq) (conjugate oxidant)

   d. $2Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2Fe^{2+}(aq)$
   Sn^{+2}(aq) (oxidant) / Sn^{4+}(aq) (conjugate reductant)
   Fe^{3+}(aq) (reductant) / Fe^{2+}(aq) (conjugate oxidant)

   e. $Cu^{2+}(aq) + Fe(s) \rightarrow Fe^{2+}(aq) + Cu(s)$
   Cu^{+2}(aq) (oxidant) / Cu(s) (conjugate reductant)
   Fe(s) (reductant) / Fe^{2+}(aq) (conjugate oxidant)

   f. $Pb(s) + Sn^{4+}(aq) \rightarrow Pb^{2+}(aq) + Sn^{2+}(aq)$
   Sn^{+4}(aq) (oxidant) / Sn^{2+}(aq) (conjugate reductant)
   Pb(s) (reductant) / Pb^{2+}(aq) (conjugate oxidant)

   g. $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$
   Cu^{+2}(aq) (oxidant) / Cu(s) (conjugate reductant)
   Mg(s) (reductant) / Mg^{2+}(aq) (conjugate oxidant)

4. Write the balanced equations for the oxidation and reduction half reactions for the balanced redox reactions that you have derived in q2. above. The first one is done for you.

$2Ag^{+}(aq) + Fe(s) \rightarrow 2Ag(s) + Fe^{2+}(aq)$

$Ag^{+}(aq)$ (oxidant) / Ag(s) (conjugate reductant)

Fe(s) (reductant) / $Fe^{2+}(aq)$ (conjugate oxidant)

Oxidation Fe(s) → $Fe^{2+}(aq)$ + 2e

Reduction-- $Ag^{+}(aq)$ + e → Ag(s)

a.

$Sn^{2+}(aq) + 2Ag^{+}(aq) \rightarrow 2Ag(s) + Sn^{4+}(aq)$

$Ag^{+}(aq)$ (oxidant) / Ag(s) (conjugate reductant)

$Sn^{2+}(aq)$ (reductant) / $Sn^{4+}(aq)$ (conjugate oxidant)

Reduction-- $Ag^{+}(aq)$ + e → Ag(s)

Oxidation-- $Sn^{2+}(aq)$ → $Sn^{4+}(aq)$ + 2e

b.*$Ca(s) + Zn^{2+}(aq) \rightarrow Zn(s) + Ca^{2+}(aq)$*

Reduction-- $Zn^{+2}(aq)$ + 2e → Zn(s)

Oxidation-- Ca(s) → $Ca^{2+}(aq)$ + 2e

c.*$Al(s) + 3Ag^{+}(aq) \rightarrow Al^{3+}(aq) + 3Ag(s)$*

Reduction-- $Ag^{+}(aq)$ + e → Ag(s)

Oxidation-- Al(s) → $Al^{3+}(aq)$ + 3e

d.*$2Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2Fe^{2+}(aq)$*

Oxidation-- $Sn^{+2}(aq)$ → $Sn^{4+}(aq)$ + 2e

Reduction-- $Fe^{3+}(aq)$ + e → $Fe^{2+}(aq)$

e.*$Cu^{2+}(aq) + Fe(s) \rightarrow Fe^{2+}(aq) + Cu(s)$*

Reduction-- $Cu^{+2}(aq)$ + 2e → Cu(s)

Oxidation-- Fe(s) → $Fe^{2+}(aq)$ + 2e

f.*$Pb(s) + Sn^{4+}(aq) \rightarrow Pb^{2+}(aq) + Sn^{2+}(aq)$*

Reduction-- $Sn^{+4}(aq)$ + 2e → $Sn^{2+}(aq)$

Oxidation-- Pb(s) → $Pb^{2+}(aq)$ + 2e

g.*$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$*

Reduction-- $Cu^{+2}(aq)$ + 2e → Cu(s)

Oxidation-- Mg(s) → $Mg^{2+}(aq)$ + 2e