Oxidation Reduction Reactions

1. What is an oxidation-reduction (or redox) reaction?

   This is a reaction in which electrons are transferred between reactants. Oxidation and Reduction must ALWAYS happen together; the only way something can gain electrons is if something is willing to give them up.

2. What do the following terms mean?

   a. Oxidation

      Oxidation is a loss of electrons

   b. Reduction

      Reduction is a gain of electrons

   c. What is a helpful way to remember this?

      A mnemonic device to help you remember:

      OIL RIG (**O**xidation **I**s **L**oss **R**eduction **I**s **G**ain)

3. What is an oxidation number?

   An oxidation number allows one to keep track of electron flow in a reaction. The method for assigning an oxidation number is based on a set of rules.

4. Assign the oxidation number for the following

   a. $KMnO_4$

      When dealing with an ionic compound, I find it easiest to break the compound up into its cation and anion.

$KMnO_4$

K^{+} $\quad$ MnO_4^{-}

charge on the monatomic potassium ion is equal to its oxidation number

sum of oxidation states must equal charge on compound

$+1$ $\qquad$ $x + 4(-2) = -1$

$x = +7$

Mn is the only unknown setting up an algebraic equation allows it to be solved for.

oxidation state of oxygen

<u>Oxidation Numbers</u>

K: $\quad$ +1
Mn: $\quad$ +7
O: $\quad$ –2 (for each oxygen)

b. $(NH_4)_2HPO_4$

$(NH_4)_2HPO_4$

NH_4^{+} $\qquad$ HPO_4^{2-}

$x + 4(+1) = +1$ $\qquad$ $+1 + x + 4(-2) = -2$

$x = -3$ $\qquad$ $x = +5$

<u>Oxidation Numbers</u>

N: $\quad$ –3
H: $\quad$ +1 (for each hydrogen)
P: $\quad$ +5
O: $\quad$ –2 (for each oxygen)

c. Fe_3O_4

In this case there is no clear cut way of separating this compound out so keep it whole…

$$Fe_3O_4$$
$$3x + 4(-2) = 0$$
$$x = {}^{8}/_{3}$$

It is important to note that it is completely valid to have an oxidation state that is a fraction, negative, etc. The reason that there is a fraction is because this composed actually contained both $Fe^{2+}$and Fe^{3+} ions. So the 8/3 is the average of the charges.

Oxidation Numbers

Fe: +8/3
O: –2 (for each oxygen)

d. $XeOF_4$

Once again this compound is not readily dissociated. Simply use the rules for assigning oxidation number and treat it as a whole.

$$XeOF_4$$
$$x - 2 + 4(-1) = 0$$
$$x = +6$$

Oxidation Numbers

Xe: +6
O: –2
F: –1 (for each fluorine)

5. What are the practical applications of oxidation numbers?

They allow us to determine what has been oxidized and what has been reduced.

6. Consider:

$$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2H_2O_{\,(g)}$$

Determine what was oxidized and what was reduced.

$$\overset{-4}{C}\overset{+1}{H_4}{}_{(g)} + 2\overset{0}{O_2}{}_{(g)} \rightarrow \overset{+4}{C}\overset{-2}{O_2}{}_{(g)} + 2\overset{+1}{H_2}\overset{-2}{O}{}_{(g)}$$

By assigning the oxidation number (shown above) you can see what happened to each element:

C: -4 → +4 Carbon became less negative (lost electrons, oxidized)
H: +1 → +1 Hydrogen didn't undergo change (unimportant)
O: 0 → -2 Oxygen gained negativity (gained electrons, reduced)

As you can see, these numbers told us a lot about what happened. If there is a reaction in which there isn't a change in oxidation number that reflects gain of electrons for one substance and loss of electrons for another it is not a redox reaction. So this method is a great way to do a quick check and determine if you are dealing with an oxidation-reduction reaction.

7. What is an

a. Oxidizing agent?

The oxidizing agent is the reactant that contains the species that gets reduced.

b. Reducing agent?

The reducing agent is the reactant that contains the species that gets oxidized.

8. Consider:

$$2\overset{+1\ \ -1}{CuCl}_{(aq)} \rightarrow \overset{+2\ \ -1}{CuCl_2}_{(aq)} + \overset{0}{Cu}_{(s)}$$

Cu: +1 → +2 Lost electrons → oxidized
Cl: -1 → -1 Nothing
Cu: +1 → 0 Gained electrons → reduced

determine the following:

a. What was oxidized?

Copper

b. What was reduced?

Copper

c. What was the reducing agent?

CuCl

d. What was the oxidizing agent?

CuCl

For this particular problem the all for answers are the same because there was only one reactant and it contains Cu which, based on oxidation numbers, is both oxidized and reduced.

9. Consider:

$$\overset{0}{Zn}_{(s)} + 2\overset{+1\ \ -1}{HCl}_{(aq)} \rightarrow \overset{+2\ \ -1}{ZnCl_2}_{(aq)} + \overset{0}{H_2}_{(g)}$$

Zn: 0 → +2 Lost electrons → oxidized
H: +1 → 0 Gained electrons→ reduced
Cl: -1 → -1 Nothing

a. What was oxidized?

Zinc

b. What was reduced?

Hydrogen

c. What was the reducing agent?

Zn

d. What was the oxidizing agent?

HCl

10. What are the steps to balancing a redox reaction using the ½ reaction method?

1. Break the reaction up into two half reactions. One for is the reduction and the other is the reduction.

2. Balance all elements in the reaction except for oxygen and hydrogen.

3. Balance the oxygen by adding H_2O.

4. Balance the hydrogen by adding H^+.

5. Balance charge by adding electrons.

6. Equalize the number of electrons coming out of one reaction with those going into the other.

7. Add up the two reactions. Make sure to cancel out where necessary.

If you asked to balance under acidic conditions – you can stop after this step. If you are asked to balance under basic conditions, you will need to proceed onto step 8.

8. Neutralize the H^+ by adding in the same number of moles of ^-OH. Make sure to add the ^-OH to both sides of the reaction.

11. Balance the following redox reaction using the ½ reaction method just described under both acidic and basic conditions.

$$MnO_{4\,(aq)} + H_2C_2O_{4\,(aq)} \rightarrow Mn^{2+}_{(aq)} + CO_{2\,(g)}$$

Step 1 – Separate reaction into two half reactions

$$MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)}$$
$$H_2C_2O_{4\,(aq)} \rightarrow CO_{2\,(g)}$$

Step 2 – Balance everything except O and H.

$$MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)}$$
$$H_2C_2O_{4\,(aq)} \rightarrow 2\,CO_{2\,(g)}$$

Step 3 – Balance O by adding H_2O

$$MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}$$
$$H_2C_2O_{4\,(aq)} \rightarrow 2\,CO_{2\,(g)}$$

Step 4 – Balance H by adding H^+

$$8H^+_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}$$
$$H_2C_2O_{4\,(aq)} \rightarrow 2\,CO_{2\,(g)} + 2H^+_{(aq)}$$

Step 5 – Balance charges by adding e^-

$$\overbrace{5e^- + \overbrace{8H^+_{(aq)} + MnO^-_{4(aq)}}^{+7}}^{+2} \rightarrow \overbrace{Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}}^{+2}$$

$$\underbrace{H_2C_2O_{4\,(aq)}}_{0} \rightarrow \underbrace{\underbrace{2\,CO_{2\,(g)} + 2H^+_{(aq)}}_{+2} + 2e^-}_{0}$$

Step 6 – Multiply each equation by a factor that will equalize the number of electrons.

$$(\,5e^- + 8H^+_{(aq)} + MnO^-_{4(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}\,)\ x2$$

$$(\,H_2C_2O_{4\,(aq)} \rightarrow 2\,CO_{2\,(g)} + 2H^+_{(aq)} + 2e^-\,)\ x5$$

$$10e^- + 16\,H^+_{(aq)} + 2\,MnO^-_{4(aq)} \rightarrow 2\,Mn^{2+}_{(aq)} + 8\,H_2O_{(l)}$$

$$5\,H_2C_2O_{4\,(aq)} \rightarrow 10\,CO_{2\,(g)} + 10\,H^+_{(aq)} + 10\,e^-$$

Step 7 – Add up the two reactions

$$\boxed{6H^+ + 2MnO_4{}^- + 5H_2C_2O_4 \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2}$$

This reaction is now balanced under acidic conditions.

In order to balance under basic conditions, it is necessary to complete step 8.

Step 8 – Add OH^- to cancel out H^+

$$\underbrace{6OH^- + 6H^+}_{\cancel{6H_2O}} + 2MnO_4^- + 5H_2C_2O_4 \rightarrow 2Mn^{2+} + \overset{2H_2O}{\cancel{8H_2O}} + 10CO_2 + 6OH^-$$

cancel out waters that are found on both sides of the reaction arrow.

$$\boxed{2MnO_4^- + 5H_2C_2O_4 \rightarrow 2Mn^{2+} + 2H_2O + 10CO_2 + 6OH^-}$$

This reaction is now balanced under basic conditions.

12. A 45.20 mL sample of sol'n containing Fe^{2+} ions is titrated with a 0.225M $KMnO_4$ sol'n. It required 23.51 mL of $KMnO_4$ sol'n to oxidize all the Fe^{2+} ions to Fe^{3+} by the following reaction

$$MnO_4^-{}_{(aq)} + Fe^{2+}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + Fe^{3+}{}_{(aq)}$$

What was the concentration of Fe^{2+} in the sol'n?

As always, determine the overall balanced first.

$$8H^+ + MnO_4^- + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$$

K^+ is a spectator ion and is disregarded in the equation. Thus the information about $KMnO_4$ is there to provide information about the MnO_4^-. Because there is a 1:1 mole ratio between $KMnO_4$ and MnO_4^- the concentration of MnO_4^- is 0.225M.

$$0.02351 \text{ L of sol'n} \; \frac{0.0225 \text{ mol } MnO_4^-}{\text{L of sol'n}} \; \frac{5 \text{ mol } Fe^{2+}}{1 \text{ mol } MnO_4^-}$$

$$= 2.645 \times 10^{-3} \text{ mol } Fe^{2+}$$

$$\frac{2.645 \times 10^{-3} \text{ mol } Fe^{2+}}{0.04520 \text{ L of sol'n}} = 0.0585 \text{ M } Fe^{2+}$$