Balancing Redox Equations Journal Notes

Note Title 20/02/2012

In previous chemistry classes you learned to balance equations following the law of Conservation of Mass. You kept the same number of atoms of each element on each side of the equation (reactants side and product side).

Many redox reactions cannot easily be balanced by just counting atoms.

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

total charge +1 ≠ total charge +2

If you simply count atoms, the equation appears to be balanced. But if you look at the charges, it isn't balanced.

There are two common methods to balance redox reactions

(A) the oxidation number method
(B) the half-reaction method

(A) Balancing equations Using Oxidation Numbers

$$Cu_{(s)} + Ag^{+1}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + Ag_{(s)}$$

	int.	final	change	
Cu	0	+2	lost 2	electrons lost by copper do not equal those gained by Ag
Ag	+1	0	gain 1 x 2	

so Ag must be multiplied by 2.

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

Balance:

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

+7 -2x4 = -8

Mn	+7	+2	5 gained	x 1	= 5	electrons now balanced
Fe	+2	+3	1 lost	x 5	= 5	

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

BUT the whole equation is not balanced. see H and O atoms. You still need to balance the other atoms in the equation.

Balance the O

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

now balance hydrogen

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

New example:

$\underset{-3 \quad +1(x3)}{NH_3} + O_2 \rightarrow \underset{+4}{NO_2} + H_2O$

				# of atoms
N	-3	+4	7 lost	1 = 7 x 4
O	0	-2	2 gained	2 = 4 x 7

One of the atoms undergoing oxidation has a subscript. O is actually O_2 so we must take this into account.

$4\ NH_3 + 7O_2 \rightarrow 4\ NO_2 + H_2O$

now balance hydrogen

$4\ NH_3 + 7O_2 \rightarrow 4\ NO_2 + 6\ H_2O$

Now try balancing practice #1.

(B) Balancing Redox Equations Using Half Reactions

This method involves breaking an equation into its two seperate components - the oxidation reaction and the reduction reaction. These seperate equations are the "half reactions".

Step 1

Break the equation into two half-reactions. If there are any spectator ions, remove them from the equation.

Step 2

Balance each half reaction seperately. First for the atoms and then for the charge. Electrons are added to one side or the other in order to balance charge.

Step 3

Make sure that the electrons lost equal electrons gained.

Step 4

Put the two half reactions back together and also any spectator ions.

$$\overset{0}{Cu}_{(s)} + \overset{+1}{Ag}\overset{+1}{NO_3}_{(aq)} \rightarrow \overset{2+}{Cu}(NO_3)_{2\,(aq)} + \overset{0}{Ag}_{(s)}$$

1. Identify which element is oxidized, reduced and any spectator ions.

oxidized - copper 0 to +2
reduced - silver +1 to 0
spectator - NO_3 ion

2. Write the two half reactions:

$$Cu \rightarrow Cu^{2+} + 2e^-$$

$$Ag^+ + 1e^- \rightarrow Ag \qquad \times 2 \text{ to balance charges}$$

atoms are balanced, charges are not.

$$Cu \rightarrow Cu^{2+} + 2e^-$$
$$2\,Ag^+ + 2e^- \rightarrow 2\,Ag$$

3. Add equations together

$$Cu + 2\,Ag^{+} \rightarrow Cu^{2+} + 2\,Ag$$

Reform compounds / Return Spectator Ions

$$Cu + 2\,AgNO_3 \rightarrow Cu(NO_3)_2 + 2\,Ag$$

Here is a reaction in an acid solution (which accounts for the H^+ ion).

$$\underset{+7\ \ -8}{MnO_4^-} + Fe^{2+} + H^+ \rightarrow Mn^{2+} + Fe^{3+} + H_2O$$

In this example the spectator ions have already been removed.

oxidation : Fe +2 to +3 $1e^-$
reduction : Mn +7 to +2 $5e^-$
spectator ions : already removed

$$Fe^{2+} \rightarrow Fe^{3+} + 1e^-$$

now the Mn reduction is a little more complex as it involves including the hydrogen and oxygen.

$$\underbrace{MnO_4^- + 8H^+ + 5e^-}_{\text{reduction}} \rightarrow Mn^{2+} + 4H_2O$$

To balance the manganese half reaction, first balance the Mn and O atoms.

Next balance H atoms and finally add enough electrons to balance the charge on both sides of the equation.

(left side $-1, +8, -5$) (right side $+2$)

$+2 = +2$

so $Fe^{2+} \rightarrow Fe^{3+} + 1e' \times 5$ $\quad 5Fe^{2+} \rightarrow 5Fe^{3+} + 5e'$

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

Add Equations together:

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

Last Example:

$HNO_3 + \overset{\circ}{Cu} + H^+ \rightarrow NO_2 + Cu^{2+} + H_2O$

Cu $\quad$ 0 to +2 oxidized

$HNO_3 + H^+ \rightarrow NO_2 + H_2O$ $\quad$ reduced

$Cu \rightarrow Cu + 2e'$

$HNO_3 + H^+ (+1e') \rightarrow NO_2 + H_2O \; (\times 2)$

balance the charges

$Cu \rightarrow Cu + 2e^-$

$2\ HNO_3 + 2H^+ + 2e^- \rightarrow 2\ NO_2 + 2H_2O$

Add Equations Together:

$2\ HNO_3 + Cu + 2H^+ \rightarrow 2\ NO_2 + Cu^{2+} + 2H_2O$

When balancing redox reactions you can use either method. You may find one easier to use.

Writing half-reactions is a skill you will need in this unit so be sure that you can write balanced half-reactions.

Now do balancing practice 2 and balancing practice 3.