Chapter 12: Oxidation and Reduction.

Oxidation-reduction (redox) reactions

At different times, ***oxidation*** and ***reduction (redox)*** have had different, but complimentary, definitions. Compare the following definitions:

Oxidation is:	Reduction is:
•Gaining oxygen	•Losing oxygen
•Losing hydrogen	•Gaining hydrogen
•Losing electrons (+ charge increases)	•Gaining electrons (+ charge decreases)

Oxidation and reduction are opposite reactions. They are also paired reactions: ***in order for one to occur, the other must also occur simultaneously***.

While the first two definitions of oxidation-reduction are correct, the most useful definition is the third - involving the gain or loss of electrons. I will use this particular definition for the rest of our discussion.

How do we know whether or not an element has gained or lost electrons? To determine which elements have been oxidized or reduced, we look at changes in the ***oxidation number*** of the element. English scientist Michael Faraday (1791 – 1867), one of the greatest pioneers in electrochemistry and electromagnetism of the 19^{th} century, developed the system of oxidation numbers to follow these kinds of reactions. The oxidation number describes the oxidation state of the element in a compound, and these numbers are assigned following a relatively simple set of rules.

Rules for assigning oxidation numbers.

1. The oxidation number of an element in its elemental form is 0 (zero).

2. The oxidation number of a simple ion is equal to the charge on the ion. Both the size and the polarity of the charge are part of the oxidation number: an ion can have a +2 oxidation number or a -2 oxidation number. The "+" and "-" signs are just as important as the number!

3. The oxidation numbers of group 1A and 2A (group 1 and group 2) elements are +1 and +2 respectively.

4. In ***compounds***, the oxidation number of hydrogen is almost always +1. The most common exception occurs when hydrogen combines with metals; in this case the oxidation number of hydrogen is typically –1.

5. In ***compounds***, the oxidation number of oxygen is almost always –2. The most common exception is in peroxides, when the oxidation number is –1. Peroxides are compounds having two oxygen atoms bonded together. For example, hydrogen peroxide is H-O-O-H. In hydrogen peroxide, each oxygen atom has a -1 oxidation number. When oxygen is bonded to fluorine, as in hypofluorous acid (HOF), the oxidation number of oxygen is 0. Oxygen-fluorine compounds are relatively rare and not too terribly important for our studies.

6. In ***compounds***, the oxidation number of fluorine is always –1. The oxidation number of other halogens (Cl, Br, I) is also –1, except when they are combined with oxygen. The oxidation number of halides (except fluorine) combined with oxygen is typically positive. For example, in ClO^-, chlorine's oxidation number is +1.

7. For a complex ion, the sum of the positive and negative oxidation numbers of all elements in the ion equals the charge on the ion.

8. For an electrically neutral compound, the sum of the positive and negative oxidation numbers of all elements in the compound equals zero.

<u>Identifying redox reactions.</u>

Now that you can assign proper oxidation numbers to the elements in substances, we can use changes in the oxidation numbers to identify oxidation and reduction reactions.

Consider the following chemical equation:

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow Zn^{+2}{}_{(aq)} + 2Cl^-{}_{(aq)} + H_{2(g)}$$

The oxidation number for elemental zinc is 0. The oxidation number for zinc ion is +2. The oxidation number for hydrogen in hydrogen chloride is +1. The oxidation number for elemental hydrogen, H_2, is 0. The oxidation number for chlorine in hydrogen chloride is –1. The oxidation number for chloride ion is –1.

Zinc has lost two electrons, and therefore developed a +2 charge. Zinc has been oxidized – its oxidation number has become more positive (from 0 to +2).

Hydrogen has gained an electron, and its positive charge has decreased. Hydrogen has been reduced – its oxidation number has become less positive (from +1 to 0).

Chloride has not changed its oxidation state. It is a spectator ion in this reaction.

Figure 12.2 may be useful in deciding if an element has been oxidized or reduced. If an elements oxidation number increases (moves towards the right), then the element is oxidized. If an elements oxidation number decreases (moves towards the left), then the element is reduced. NOTE: an element doesn't have to become positive or negative for oxidation or reduction to occur. Instead, the element has to become *more positive* or *more negative*. A change from -3 to -1 is still oxidation, while a change from +3 to +1 is still reduction.

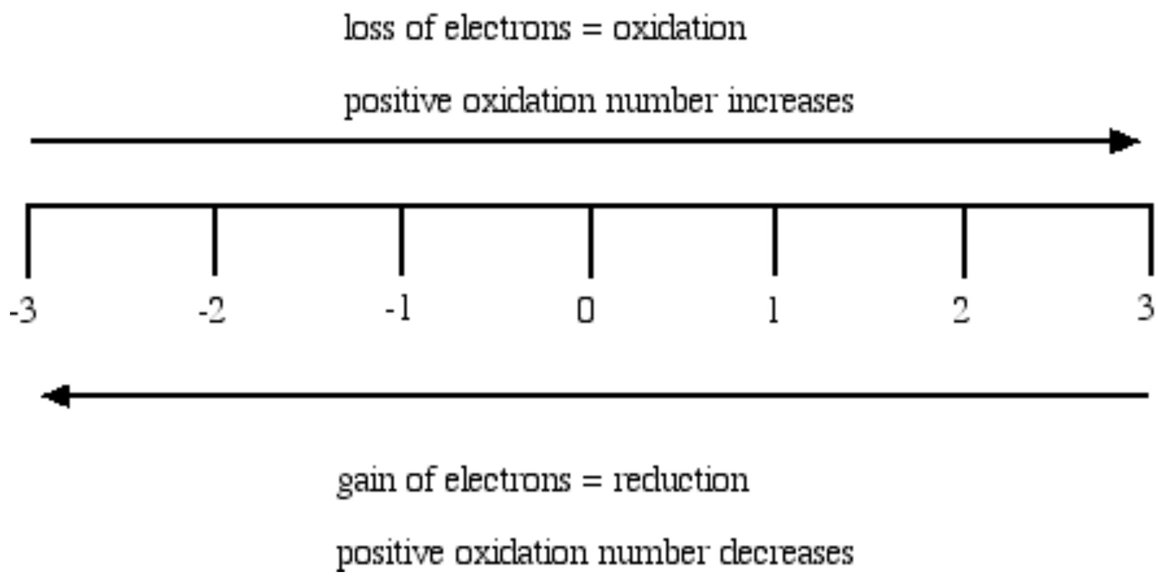

Figure 12.2. Graphic description of oxidation and reduction.

Substances that cause changes in the oxidation state are called ***oxidizing agents*** or ***reducing agents***. An oxidizing agent *causes* oxidation number to occur. How does the oxidizing agent cause oxidation?

To increase the positive oxidation number of an element, the oxidizing agent must take one or more electrons from the element. As the element being oxidized loses electron(s), its oxidation number becomes more positive. However, the electrons don't disappear! The oxidizing agent has taken these electrons, and therefore the oxidizing agent becomes more negative – *it is reduced*!

In our reaction above, hydrogen in hydrogen chloride takes an electron from the zinc metal. The zinc metal becomes more positive; it is oxidized. By taking the electron from the zinc metal, the hydrogen in hydrogen chloride becomes less positive; it is reduced. Hydrogen chloride is the oxidizing agent, because it contains the element that causes oxidation to occur.

Similarly, the zinc metal donates electrons to the hydrogen in hydrogen chloride, causing the oxidation state of hydrogen to decrease from +1 to 0. By providing the electrons necessary to reduce the oxidation number of hydrogen, the oxidation number of zinc increases from 0 to +2; *zinc is oxidized*. While being oxidized, zinc reduced the oxidation number of hydrogen. Therefore, zinc is a reducing agent.

Oxidizing agents are substances containing the element(s) that accept electrons, allowing another element(s) to be oxidized. By accepting electrons, the element(s) in the oxidizing agent are reduced.

Reducing agents are substances containing the element(s) that donate electrons, allowing another element(s) to be reduced. By donating electrons, the element(s) in the reducing agent are oxidized.

As you can see, there is a great deal of symmetry between oxidizing and reducing agents, and between oxidation and reduction. Whenever one process happens, the other process **MUST** also happen, because we are transferring electrons from one material to another material. Electrons must come from someplace, and must go someplace; electrons cannot simply appear and disappear.

There are two general considerations to keep in mind when discussing oxidizing/reducing agents. First, in most cases there will only be one element in each agent that is oxidized or reduced. In our reaction, only hydrogen in hydrogen chloride was reduced – the chloride wasn't changed. Second, the oxidizing or reducing agent is the ***compound*** containing the element that is oxidized or reduced. The reducing agent is hydrogen chloride, not just hydrogen ion. Similarly, the oxidizing agent is the ***compound*** containing the element that is reduced. Oxidizing and reducing agents are **ALWAYS** reactants, **NEVER** products!

Balancing redox reactions.

We can use our knowledge of redox reactions to help us balance chemical reactions. The simplest process to use is the ***half-reaction method***. Consider the following chemical reaction (unbalanced):

$$MnO_4^- + C_2O_4^{-2} \rightarrow Mn^{+2} + CO_2$$

First, we assign oxidation numbers to all elements shown in the reaction:

$$\overset{+7\ -2}{MnO_4^-} + \overset{+3\ -2}{C_2O_4^{-2}} \rightarrow \overset{+2}{Mn^{+2}} + \overset{+4\ -2}{CO_2}$$

Comparing oxidation numbers, we see that manganese goes from +7 to +2, and has been reduced, while carbon goes from +3 to +4 and has been oxidized. Oxygen doesn't change its oxidation state. The oxidizing agent is permanganate, and the reducing agent is oxalate.

The half-reaction method involves balancing the oxidation reaction as if it were an isolated reaction. Then the reduction half-reaction is balanced as if it were an isolated reaction. Finally, the two half-reactions are combined.

The oxidation half-reaction is:

$$C_2O_4^{-2} \rightarrow CO_2$$

The first step is to balance elements, other than oxygen and hydrogen, using the normal methods of balancing chemical equations. We have two carbons in oxalate, so we need a 2 in front of carbon dioxide to balance the carbons:

$$C_2O_4^{-2} \rightarrow 2CO_2$$

By comparing reactants and products, we see that we have the same number of carbon and oxygen atoms on each side of the reaction. There aren't any other elements present, so we don't need to change any more coefficients.

Charge balance requires that the net charge of reactants and products must be equal. We equalize the charges by adding 2 electrons to the products side:

$$C_2O_4^{-2} \rightarrow 2CO_2 + 2e^-$$

All elements are identical on both sides, the number of atoms of each element is the same on both sides, and the total charge is the same on both sides: -2 for oxalate and -2 for the two electrons. This is a balanced oxidation half-reaction.

The reduction half-reaction is:

$$MnO_4^- \rightarrow Mn^{+2}$$

The manganese atoms are balanced, but what are we to do about oxygen? To balance oxygen, we add water molecules. Four oxygen atoms in permanganate require four water molecules as products:

$$MnO_4^- \rightarrow Mn^{+2} + 4H_2O$$

Adding water molecules balanced oxygen, but now we have hydrogen atoms. We balance 8 hydrogen atoms from 4 water molecules by adding 8 hydrogen ions to the reaction:

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

Now we have the same number and kinds of atoms on both sides. However, charge balance requires that the net charge of reactants and products must be equal. We have a net electrical charge of +7 for the reactants, and a net electrical charge of +2 for the products. We add 5 electrons to the reactant side of the equation:

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

This is a balanced reduction half-reaction. (NOTE: in an oxidation half-reaction, electrons are produced, while in a reduction half-reaction, electrons are consumed.)

We are ready to combine the two half-reactions. To combine these reactions, the same number of electrons must be produced as are consumed. When different numbers of electrons are produced and consumed, you need to find the least common multiple. Since 5 electrons are consumed in the reduction reaction, and 2 electrons are produced in the oxidation reaction, the least common multiple is 10. We need to multiply the reduction reaction by 2, and the oxidation reaction by 5:

$$16H^+ + 2MnO_4^- + 10e^- \rightarrow 2Mn^{+2} + 8H_2O$$

$$5C_2O_4^{-2} \rightarrow 10CO_2 + 10e^-$$

Now that we have the same number of electrons produced and consumed, and we can add the reactions:

$$16H^+ + 2MnO_4^- + 10e^- + 5C_2O_4^{-2} \rightarrow 2Mn^{+2} + 8H_2O + 10CO_2 + 10e^-$$

We eliminate substances that are identical on both sides of the equation:

$$16H^+ + 2MnO_4^- + 5C_2O_4^{-2} \rightarrow 2Mn^{+2} + 8H_2O + 10CO_2$$

Finally, we check the equation to insure that mass and charge balance have been achieved.

Reactants	Products
16 H	16 H
2 Mn	2 Mn
28 O	28 O
10 C	10 C
+4	+4

This is an example of balancing a redox reaction in acid solution, by the half-reaction method. We can also balance the reaction in basic solution. The initial work is the same as shown above. Once the reaction is balanced in acid solution, we neutralize hydrogen ion by adding an equal number of hydroxide ions to both sides of the reaction:

$$16OH^- + 16H^+ + 2MnO_4^- + 5C_2O_4^{-2} \rightarrow 2Mn^{+2} + 8H_2O + 10CO_2 + 16OH^-$$

Hydroxide ion combines with hydrogen ion, forming water molecules, and we eliminate equal numbers of water molecules from both sides of the equation:

$$16H_2O + 2MnO_4^- + 5C_2O_4^{-2} \rightarrow 2Mn^{+2} + 8H_2O + 10CO_2 + 16\ OH^-$$

$$8H_2O + 2MnO_4^- + 5C_2O_4^{-2} \rightarrow 2Mn^{+2} + 10CO_2 + 16OH^-$$

Finally, we check to insure mass and charge balance.

Reactants	Products
16 H	16 H
2 Mn	2 Mn
36 O	36 O
10 C	10 C
-12	-12

Vocabulary. The following terms are defined and explained in the text. Make sure that you are familiar with the meanings of the terms as used in chemistry. Understand that you may have been given incomplete or mistaken meanings for these terms in earlier courses. The meanings given in the text are correct and proper.

Oxidation	**Reduction**	**Oxidation number**
Oxidizing agent	**Reducing agent**	

Assign correct oxidation numbers to all elements in the following substances.

1. H_2S
2. BaO
3. P_2O_5
4. $Cr_2O_7^{-2}$
5. CH_2O
6. HBr
7. $Fe(NO_3)_3$
8. CO
9. $HOCl$
10. H_2SO_4

In the following reactions, identify the oxidizing agent and reducing agent.

1. $Fe + Ni(NO_3)_2 \rightarrow Fe(NO_3)_2 + Ni$
2. $Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2H_2O + 2NO_2$
3. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
4. $C_2H_4 + H_2 \rightarrow C_2H_6$
5. $H_2SO_4 + 2NaOH \rightarrow 2Na^+ + SO_4^{-2} + 2H_2O$

Use the half-reaction method and balance the following equations under the listed conditions.

1. $CN^- + MnO_4^- \rightarrow CNO^- + MnO_2$ (basic solution)
2. $Cr_2O_7^{-2} + I^- \rightarrow Cr^{+3} + IO_3^-$ (acidic solution)
3. $MnO_4^- + CH_3OH \rightarrow Mn^{+2} + HCO_2H$ (acidic solution)
4. $As + ClO_3^- \rightarrow H_3AsO_3 + HClO$ (acidic solution)
5. $H_2O_2 + ClO_2 \rightarrow ClO_2^- + O_2$ (basic solution)

Answers:

Assigning oxidation numbers.

1. The oxidation number of hydrogen is +1. Since there are 2 hydrogen atoms in the compound, and no net electrical charge is shown, the oxidation number of sulfur must be –2:

$$\overset{+1}{H_2}\overset{-2}{S}$$

2. The oxidation number of oxygen is –2. With no net electrical charge shown, the oxidation number of barium is +2. (Note: barium is a group IIA (2) element, and by rule 3 would have an oxidation number of +2.)

3. The oxidation number of oxygen is –2. There is no net electrical charge shown. Five oxygen atoms produce a total negative oxidation state of –10. There must be a total of +10 oxidation numbers on phosphorous. With two phosphorous atoms, we assign an oxidation number of +5 to each phosphorous.

4. The oxidation number of oxygen is –2. For seven oxygen atoms, this produces a total of –14. There is a -2 charge shown for the ion; therefore two of the –14 oxidation numbers are assigned to the charge. This leaves –12, which must be balanced by the positive oxidation state of chromium. With two chromium atoms in the ion, we assign each an oxidation number of +6.

5. The oxidation number of oxygen is –2. The oxidation number of hydrogen is +1. Since the total positive oxidation state (2 x +1) exactly balances the total negative oxidation state, the oxidation state of carbon must be zero (0).

   NOTE: while the oxidation state of elements in their elemental form **must** be zero, it is also possible for an element to have a zero oxidation state when combined in a compound.

6. The oxidation number of hydrogen is +1, while that of bromine is –1.

7. It is important (and easiest) if you recognize this as an ionic compound composed of the simple ion Fe^{+3}, and the polyatomic anion NO_3^-.

   The oxidation number of the iron is equal to its charge, +3. In the polyatomic anion, the oxidation number of each oxygen atom is –2, for a total of –6. One of these negative oxidation numbers is the charge on the ion, -1. Therefore, there are five negative oxidation numbers to be balanced by the positive oxidation number of nitrogen, which must have a value of +5.

8. The oxidation state of oxygen is –2, therefore carbon must be +2.

9. The oxidation state of hydrogen is +1. The oxidation state of oxygen is –2, therefore the oxidation state of chlorine must be +1.

10. The oxidation state of hydrogen is +1. The oxidation state of oxygen is –2. Four oxygen atoms result in a total of –8, while two hydrogen atoms produce a total of +2. With no net charge shown, the oxidation state of sulfur must be +6.

Identifying oxidizing and reducing agents.

1. Elemental iron has an oxidation number of 0. The oxidation number of iron in iron(II) nitrate is +2. The oxidation number of nickel in nickel(II) nitrate is +2. The oxidation number of elemental nickel is 0. The oxidation numbers of nitrogen and oxygen don't change; nitrogen is +5 and oxygen is -2. Therefore, iron is oxidized (it is the reducing agent), while nickel ion is reduced (nickel(II) nitrate is the oxidizing agent).

2. Elemental copper has an oxidation number of 0. The copper in copper(II) nitrate has an oxidation number of +2. The oxidation number of hydrogen in all compounds shown is +1 (it did not change). The oxidation number of oxygen in all compounds shown is –2 (it also has not changed). The oxidation number of nitrogen in the nitric acid is +5. The oxidation number of nitrogen in the copper(II) nitrate is also +5. However, the oxidation number of nitrogen in nitrogen dioxide is +4. Therefore, copper has lost electrons and been oxidized (it is the reducing agent). The nitrogen in nitric acid has gained electrons and has been reduced (nitric acid is oxidizing agent).

3. The oxidation number of carbon in methane (CH_4) is –4, while its oxidation number in carbon dioxide is +4. The oxidation number of hydrogen in all compounds is +1. The oxidation number of oxygen in its elemental form is 0, while its value in carbon dioxide and water is –2. Therefore, carbon has been oxidized, and methane is the reducing agent. Oxygen has been reduced, and it is the oxidation agent.

4. The oxidation number of carbon in ethene (C_2H_4) is –2, while in ethane (C_2H_6) carbon has an oxidation number of –3. The oxidation number of hydrogen in ethene and ethane is +1. The oxidation number of elemental hydrogen is 0. Therefore, carbon in ethene has been reduced, and ethene is the oxidizing agent. Elemental hydrogen has been oxidized, and it is the reducing agent.

5. No element changed its oxidation number in this reaction. Nothing has been oxidized, and nothing has been reduced. This is not a redox reaction.

Balancing reactions by half reaction method.

1. $CN^- + MnO_4^- \rightarrow CNO^- + MnO_2$ (basic solution)

   a. $CN^- \rightarrow CNO^-$ oxidation reaction

   b. $H_2O + CN^- \rightarrow CNO^-$ balance O with water

   c. $H_2O + CN^- \rightarrow CNO^- + 2H^+$ balance H with H^+

   d. $H_2O + CN^- \rightarrow CNO^- + 2H^+ + 2e^-$ balance charge with e^-

   e. $MnO_4^- \rightarrow MnO_2$ reduction reaction

   f. $MnO_4^- \rightarrow MnO_2 + 2H_2O$ balance O with water

   g. $4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O$ balance H with H^+

   h. $4H^+ + MnO_4^- + 3e^- \rightarrow MnO_2 + 2H_2O$ balance charge with e^-

   i. multiply "d" x 3: multiply "h" x 2

      $3H_2O + 3CN^- \rightarrow 3CNO^- + 6H^+ + 6e^-$
      $8H^+ + 2MnO_4^- + 6e^- \rightarrow 2MnO_2 + 4H_2O$

   j. add two equations together

      $3H_2O + 3CN^- + 8H^+ + 2MnO_4^- + 6e^-$
      $\rightarrow 3CNO^- + 6H^+ + 6e^- + 2MnO_2 + 4H_2O$

   k. cancel common substances on both sides

      $3CN^- + 2H^+ + 2MnO_4^- \rightarrow 3CNO^- + 2MnO_2 + H_2O$

   l. add OH^- to both sides to neutralize H^+

      $3CN^- + 2H^+ + 2OH^- + 2MnO_4^-$
      $\rightarrow 3CNO^- + 2MnO_2 + H_2O + 2OH^-$

   m. combine H+ with OH-, cancel out waters.

      $3CN^- + H_2O + 2MnO_4^- \rightarrow 3CNO^- + 2MnO_2 + 2OH^-$

   n. check reactants/products for mass/charge balance

	Reactants	Products
C	3	3
N	3	3
H	2	2
O	9	9
Mn	2	2
Charge	-5	-5

2. $Cr_2O_7^{-2} + I^- \rightarrow Cr^{+3} + IO_3^-$ (acidic solution)

a. $Cr_2O_7^{-2} \rightarrow Cr^{+3}$ reduction reaction

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

c. $Cr_2O_7^{-2} \rightarrow 2Cr^{+3} + 7H_2O$ balance O with water

d. $Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O$ balance H with H^+

e. $Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$ balance charge with e^-

f. $I^- \rightarrow IO_3^-$ oxidation reaction

g. $I^- + 3H_2O \rightarrow IO_3^-$ balance O with water

h. $I^- + 3H_2O \rightarrow IO_3^- + 6H^+$ balance H with H^+

i. $I^- + 3H_2O \rightarrow IO_3^- + 6H^+ + 6e^-$ balance charge with e^-

j. Add “e” and “i” together

$Cr_2O_7^{-2} + 14H^+ + 6e^- + I^- + 3H_2O$
$\rightarrow 2Cr^{+3} + 7H_2O + IO_3^- + 6H^+ + 6e^-$

k. cancel common substances on both sides of equation

$Cr_2O_7^{-2} + 8H^+ + I^- \rightarrow 2Cr^{+3} + 4H_2O + IO_3^-$

l. check reactants/products for mass/charge balance

	Reactants	Products
Cr	2	2
O	7	7
H	8	8
I	1	1
Charge	+5	+5

3. $MnO_4^- + CH_3OH \rightarrow Mn^{+2} + HCO_2H$ (acidic solution)

a, $MnO_4^- \rightarrow Mn^{+2}$ reduction reaction

b. $MnO_4^- \rightarrow Mn^{+2} + 4H_2O$ balance O with water

c. $8H^+ + MnO_4^- \rightarrow Mn^{+2} + 4H_2O$ balance H with H^+

d. $8H^+ + 5e^- + MnO_4^- \rightarrow Mn^{+2} + 4H_2O$ balance charge with e^-

e. $CH_3OH \rightarrow HCO_2H$ oxidation reaction

f. $H_2O + CH_3OH \rightarrow HCO_2H$ balance O with water

g. $H_2O + CH_3OH \rightarrow HCO_2H + 4H^+$ balance H with H^+

h. $H_2O + CH_3OH \rightarrow HCO_2H + 4H^+ + 4e^-$ balance charge with e^-

i. multiply “d” by 4, multiply “h” by 5

$32H^+ + 20e^- + 4MnO_4^- \rightarrow 4Mn^{+2} + 16H_2O$
$5H_2O + 5CH_3OH \rightarrow 5HCO_2H + 20H^+ + 20e^-$

j. add two equations together

$32H^+ + 20e^- + 4MnO_4^- + 5H_2O + 5CH_3OH$
$\rightarrow 4Mn^{+2} + 16H_2O + 5HCO_2H + 20H^+ + 20e^-$

k. cancel out common substances on both sides of equation

$12H^+ + 4MnO_4^- + 5CH_3OH \rightarrow 4Mn^{+2} + 11H_2O + 5HCO_2H$

l. check reactants/products for mass/charge balance

	Reactants	Products
H	32	32
Mn	4	4
O	21	21
C	5	5
Charge	+8	+8

4. $As + ClO_3^- \rightarrow H_3AsO_3 + HClO$ (acidic solution)

a. $As \rightarrow H_3AsO_3$ oxidation reaction

b. $3H_2O + As \rightarrow H_3AsO_3$ balance O with water

c. $3H_2O + As \rightarrow H_3AsO_3 + 3H^+$ balance H with H^+

d. $3H_2O + As \rightarrow H_3AsO_3 + 3H^+ + 3e^-$ balance charge with e^-

e. $ClO_3^- \rightarrow HClO$ reduction reaction

f. $ClO_3^- \rightarrow HClO + 2H_2O$ balance O with water

g. $5H^+ + ClO_3^- \rightarrow HClO + 2H_2O$ balance H with H^+

h. $5H^+ + 4e^- + ClO_3^- \rightarrow HClO + 2H_2O$ balance charge with e^-

i. multiply “d” by 4, multiply “h” by 3

$12H_2O + 4As \rightarrow 4H_3AsO_3 + 12H^+ + 12e^-$
$15H^+ + 12e^- + 3ClO_3^- \rightarrow 3HClO + 6H_2O$

j. add two equations together

$12H_2O + 4As + 15H^+ + 12e^- + 3ClO_3^-$
$\rightarrow 4H_3AsO_3 + 12H^+ + 12e^- + 3HClO + 6H_2O$

k. cancel out common substances on both sides of equation

$6H_2O + 4As + 3H^+ + 3ClO_3^- \rightarrow 4H_3AsO_3 + 3HClO$

l. check reactants/products for mass/charge balance

	Reactants	Products
H	15	15
O	15	15
As	4	4
Cl	3	3
Charge	0	0

5. $H_2O_2 + ClO_2 \rightarrow ClO_2^- + O_2$ (basic solution)

a, $ClO_2 \rightarrow ClO_2^-$ reduction equation

b. $e^- + ClO_2 \rightarrow ClO_2^-$ balance charge with e^-

c. $H_2O_2 \rightarrow O_2$ oxidation reaction

d. $H_2O_2 \rightarrow O_2 + 2H^+$ balance H with H^+

e. $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$ balance charge with e^-

f. multiply "b" by 2 and add to "e"

$2e^- + 2ClO_2 + H_2O_2 \rightarrow 2ClO_2^- + O_2 + 2H^+ + 2e^-$

g. cancel out common substances on both sides of equation

$2ClO_2 + H_2O_2 \rightarrow 2ClO_2^- + O_2 + 2H^+$

h. add OH^- to both sides to neutralize H^+

$2ClO_2 + H_2O_2 + 2OH^- \rightarrow 2ClO_2^- + O_2 + 2H^+ + 2OH^-$

$2ClO_2 + H_2O_2 + 2OH^- \rightarrow 2ClO_2^- + O_2 + 2H_2O$

i. check reactants/products for mass/charge balance

	Reactants	Products
Cl	2	2
O	8	8
H	4	4
Charge	-2	-2