ELECTROCHEMISTRY

A branch of chemistry which deals with the chemical changes produced by an electric current and with the production of electricity by means of a chemical reaction.

Chemical reactions in which electrons are transferred from one substance to another are called oxidation-reduction or “redox” reactions after the two processes which take place in the reactions (oxidation and reduction).

OXIDATION

The term “oxidation” originally meant a reaction in which oxygen combines with another substance.

Oxidation “half-reactions” are chemical reactions in which electrons are lost by a substance.

When a substance is *oxidized,* it loses electrons and its oxidation number or charge becomes more positive.

REDUCTION

Reduction "half-reactions" are chemical reactions in which electrons are gained by a substance.

When a substance is *reduced,* it gains electrons and its oxidation number or charge becomes more negative.

Oxidation and reduction "half-reactions" are simultaneous processes. One can not occur without the other.

A substance can not lose electrons unless another substance gains those very electrons.

During oxidation-reductions reaction, the *oxidation number or charge* on the substance changes.

OXIDATION NUMBER (STATE)

The oxidation number or oxidation state of an element is the number of electrons lost, gained, or shared by that substance.

Oxidation

- The oxidation number increases
- Electrons are lost making the oxidation number more positive

+7

+6

+5

+4

+3

+2

+1

0

-1

-2

-3

-4

Reduction

- The oxidation number decreases
- Electrons are gained making the oxidation number more negative

FINDING OXIDATION NUMBERS

â The oxidation number or state of all uncombined elements is “zero.”

ã The oxidation number of all IA (sodium, potassium, lithium, etc.) particles is always +1 unless they are found as atoms. IIA elements always form the +2 oxidation state when combined.

ä Hydrogen always has a +1 oxidation number in chemical compounds unless it has combined with a metal when hydrogen becomes -1.

å Oxygen always has a -2 oxidation number except when oxygen has combined with a fluorine or itself (peroxide).

æ The sum of the oxidation numbers of the elements in a compound equals “zero.”

ç The sum of the oxidation numbers of the elements in an ion equals the charge on that ion.

è Elements lose and gain electrons according to their electronegativity.

é Elements with high electronegativity (nonmetals) gains electrons to form negative oxidation numbers.

ê Elements with low electronegativity (metals) lose electrons to form positive oxidation numbers.

ë Combined fluorine always has a -1 oxidation number since it has the highest electronegativity of all elements (4.0).

What is the oxidation number of carbon in carbon dioxide, CO_2?

$$\overset{x\ \ -2}{CO_2}$$

$$x + 2(-2) = 0$$

$$x = +4$$

What is the oxidation number of oxygen in OF_2?

$$\overset{x\ \ -1}{OF_2}$$

$$x + 2(-1) = 0$$

$$x = +2$$

What is the oxidation number of nitrogen in NH_3?

$$\begin{array}{c} x \quad +1 \\ NH_3 \end{array}$$

$$x + 3(+1) = 0$$

$$x = -3$$

What is the oxidation number of Mn in $KMnO_4$?

$$\begin{array}{c} +1 \;\; x \;\; -2 \\ KMnO_4 \end{array}$$

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

$$x = +7$$

Determine the oxidation number of Cr in $K_2Cr_2O_7$?

$$\overset{+1}{K_2}\overset{2x}{Cr_2}\overset{-2}{O_7}$$

$$2(+1) + 2x + 7(-2) = 0$$

$$2x = 12$$

$$x = +6$$

What is the oxidation number of nitrogen in the nitrate ion, NO_3^{-1}?

$$\overset{x}{N}\overset{-2}{O_3}{}^{-1}$$

$$x + 3(-2) = -1$$

$$x = +5$$

Find the oxidation number of chlorine in each of the following substances.

Cl_2 $\quad$ HCl $\quad$ $NaClO_3$ $\quad$ Cl_2O_7

$$\overset{x}{Cl_2}$$
$$2x = 0$$
$$x = 0$$

$$\overset{+1}{H}\overset{x}{Cl}$$
$$1 + x = 0$$
$$x = -1$$

$$\overset{+1}{Na}\overset{x}{Cl}\overset{-2}{O_3}$$

$$\overset{2x}{Cl_2}\overset{-2}{O_7}$$

$$1 + x + 3(-2) = 02x + 7(-2) = 0$$

$$x = +5 \qquad x = +7$$

What is the oxidation number of xenon, Xe, in $XeOF_2$?

$$\overset{x}{Xe}\ \overset{-2}{O}\ \overset{-1}{F_2}$$

$$x + (-2) + 2(-1) = 0$$

$$x = +4$$

What is the oxidation number of hydrogen in boron hydride, BH_3?

$$\overset{+3}{B}\overset{x}{H_3}$$

$$3 + 3x = 0$$

$$x = -1$$

What is the oxidation number of sulfur in $(NH_4)_2SO_4$?

$$\overset{+1}{(NH_4)_2}\overset{x}{S}\overset{-2}{O_4}$$

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

$$x = +6$$

Determine the oxidation number of oxygen in hydrogen peroxide, H_2O_2.

$$\overset{+1}{H_2}\overset{x}{O_2}$$

$$2(+1) + 2x = 0$$

$$x = -1$$

In the reaction:

$$2NO + O_2 \ 6 \ 2NO_2$$

What is the change in the oxidation number of nitrogen?

$$\overset{x\ -2}{NO}$$

$$x + (-2) = 0$$

$$x = +2$$

$$\overset{x\ -2}{NO_2}$$

$$x + 2(-2) = 0$$

$$x = +4$$

The oxidation number changes from +2 to +4.

Is this change oxidation or reduction?

Answer: Oxidation because the oxidation number increases (becomes more positive).

Find the change in the oxidation number of manganese in the following reaction.

$16HCl + 2KMnO_4$ 6
$2MnCl_2 + 2KCl + 5Cl_2 + 8H_2O$

$$\overset{+1}{K}\overset{x}{Mn}\overset{-2}{O_4} \qquad \overset{x}{Mn}\overset{-1}{Cl_2}$$

$$1 + x + 4(-2) = 0 \qquad x + 2(-1) = 0$$

$$x = +7 \qquad x = +2$$

The oxidation number changes from +7 to +2.

Is this change oxidation or reduction?

Answer: Reduction because the oxidation number decreases (becomes more negative).

What is the change in the oxidation number of nitrogen in the following unbalanced equation?

$$HNO_3 + HCl \rightarrow NO + Cl_2 + H_2O$$

$$\overset{+1}{H}\overset{x}{N}\overset{-2}{O_3} \qquad\qquad \overset{x}{N}\overset{-2}{O}$$

$$+1 + x + 3(-2) = 0 \qquad x + 1(-2) = 0$$

$$x = +5 \qquad\qquad x = +2$$

The change in the oxidation number of nitrogen is from +5 to +2.

Is this change oxidation or reduction?

Answer: Reduction because the oxidation number decreases (becomes more negative).

Finding Changes in Oxidation States

â The presence of an uncombined element in a chemical reaction almost always signifies a change in oxidation state.

ã Elements with multiple oxidation states will most likely undergo the change in a chemical reaction. Likewise, hydrogen, oxygen, the IA and IIA elements usually do not change.

ä Since elements react in simple, whole number ratios, a change in the formula or ratio <u>may</u> be a result of a change in the oxidation state.

å The formulas of polyatomic ions are determined by the oxidation states of the elements present in the ion. If the formula of the ion has not changed, the oxidation state probably has not changed.

æ There are always at least two changes, one oxidation and one reduction.

Determine the two elements which change oxidation numbers in the following reaction. What are the changes? Which is oxidation and which is reduction?

$$Cu + H_2SO_4 \ 6 \ CuSO_4 + SO_2 + H_2O$$

$$\overset{0}{Cu} \quad 6 \quad \overset{x\ -2}{CuSO_4}$$

$$x + 1(-2) = 0$$

$$x = +2$$

$$\overset{+1\ x\ -2}{H_2SO_4} \quad 6 \quad \overset{x\ -2}{SO_2}$$

$$2(+1) + x + 4(-2) = 0 \qquad x + 2(-2) = 0$$

$$x = +6 \qquad x = +4$$

The copper is oxidized (0 to +2)

The sulfur is reduced (+6 to +4)

Given the unbalanced equation:

$$NH_3 + NaBrO \text{ 6 } NaBr + H_2O + N_2$$

Which element is oxidized and which element is reduced?

$$\overset{x\ +1}{NH_3} \qquad \text{6} \qquad \overset{0}{N_2}$$

$$x + 3(+1) = 0$$

$$x = -3$$

$$\overset{+1\ x\ -2}{NaBrO} \qquad \text{6} \qquad \overset{+1\ x}{NaBr}$$

$$+1 + x + 1(-2) = 0 \qquad\qquad +1 + x = 0$$

$$x = +1 \qquad\qquad x = -1$$

Nitrogen is oxidized (-3 to 0)
Bromine is reduced (+1 to -1)

Given the unbalanced equation:

$$FeSO_4 + H_2SO_4 + K_2Cr_2O_7 \text{ 6}$$

$$Fe_2(SO_4)_3 + K_2SO_4 + Cr_2(SO_4)_3 + H_2O$$

Which element is oxidized and which element is reduced?

$$\overset{x}{Fe}\overset{-2}{SO_4}$$
$$x - 2 = 0$$
$$x = +2$$

6

$$\overset{x}{Fe_2}(\overset{-2}{SO_4})_3$$
$$2x + 3(-2) = 0$$
$$x = +3$$

$$\overset{+1}{K_2}\overset{x}{Cr_2}\overset{-2}{O_7}$$
$$2(+1) + 2x + 7(-2) = 0$$
$$x = +6$$

6

$$\overset{x}{Cr_2}(\overset{-2}{SO_4})_3$$
$$2x + 3(-2) = 0$$
$$x = +3$$

Iron is oxidized (+2 to +3)
Chromium is reduced (+6 to +3)

Given the unbalanced equation:

$$As_2O_3 + HBrO_3 + H_2O \; 6 \; H_3AsO_4 + HBr$$

Which element is oxidized and which element is reduced?

$$\overset{x\;\;-2}{As_2O_3} \quad 6 \quad \overset{+1\;\;x\;\;-2}{H_3AsO_4}$$

$$2x + 3(-2) = 0 \qquad 3(+1) + x + 4(-2) = 0$$

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

$$\overset{+1\;\;x\;\;-2}{HBrO_3} \quad 6 \quad \overset{+1\;\;x}{HBr}$$

$$+1 + x + 3(-2) = 0 \qquad +1 + x = 0$$

$$x = +5 \qquad x = -1$$

Arsenic is oxidized (+3 to +5)
Bromine is reduced (+5 to -1)

BALANCING OXIDATION-REDUCTION HALF-REACTIONS

Oxidation-reduction half-reactions are balanced when:

â The equation contains the same number of each element on both sides of the equation.

ã The sum of the charges on both sides of the equation is the same.

ä If the sum of the charges is not the same on both sides, always add electrons to the side of the half-reaction with the more positive charge.

$$As^{3+} \ 6 \ As^{5+}$$

$$Br^{1+} \ 6 \ Br^{1-}$$

Balance the each of following half-reactions by adding the correct number of electrons to the appropriate side.

â Fe^{0} 6 Fe^{3+}

ã Mn^{7+} 6 Mn^{2+}

ä $Cl_2^{\,0}$ 6 $2Cl^{-}$

å $NO_3^{-} + 4H^{+}$ 6 $NO + 2H_2O$

æ $SO_2 + 2H_2O$ 6 $SO_4^{2-} + 4H^{+}$

â Fe^{0} 6 $Fe^{3+} + 3e^{-}$

ã $Mn^{7+} + 5e^{-}$ 6 Mn^{2+}

ä $Cl_2^{\,0} + 2e^{-}$ 6 $2Cl^{-}$

å $NO_3^{-} + 4H^{+} + 3e^{-}$ 6 $NO + 2H_2O$

æ $SO_2 + 2H_2O$ 6 $SO_4^{2-} + 4H^{+} + 2e^{-}$

Which of the following half-reactions correctly represents oxidation?

(a) $Cl_2^0 + 2e^-$ 6 $2Cl^-$

(b) $2Cl^- + 2e^-$ 6 Cl_2^0

(c) Cl_2^0 6 $2Cl^- + 2e^-$

(d) $2Cl^-$ 6 $Cl_2^0 + 2e^-$

Which of the following reactions correctly represents reduction?

(a) $Al^{3+} + 3e^-$ 6 $3Al^0$

(b) $3Al^0$ 6 $Al^{3+} + 3e^-$

(c) $Al^{3+} + 3e^-$ 6 Al^0

(d) Al^0 6 $Al^{3+} + 3e^-$

BALANCING REDOX REACTIONS

Oxidation-reduction (redox) reactions are balanced when:

â A balanced redox equation contains the same number of each element on both sides of the equation.

ã The sum of the charges on both sides of the equation is the same.

ä The number of electrons lost equals the number of electrons gained.

Which of the following oxidation-reduction reactions are correctly balanced?

â $Zn(s) + H^{+}(aq)$ 6 $Zn^{2+}(aq) + H_2(g)$

ã $2H^{+}(g) + CuO(s)$ 6 $Cu^{0}(s) + H_2O$(R)

ä $5I^{-}(aq) + IO_3^{-}(aq) + 6H^{+}(aq)$6
$3I_2(s) + 3H_2O$(R)

å $2Na(s) + 2H_2O$(R) 6
$2Na^{+}(aq) + 2OH^{-}(aq) + H_2(g)$

æ $Cr_2O_7^{2-} + 3Sn^{2+} + 8H^{+}$ 6
$Cr^{3+} + 7H_2O + Sn^{4+}$

STEPS NEEDED TO BALANCE REDOX REACTIONS

â Determine the two changes in oxidation number, one change must be oxidation and the other change must be reduction.

ã Balance each half-reaction for mass (number of atoms must be the same on both sides of the equation)

ä Add electrons to the more positive side of each half-reaction to balance it for charge (the sum of the charges must be the same on both sides of the equation)

å Using the lowest common denominator between the electrons lost and the electrons gained, multiply each half-reaction by the quantity which will make the number of electrons lost equal to the number of electrons gained.

æ Add the two half-reactions together to obtain the net ionic equation.

ç Transfer the coefficients from the net ionic equation to the formula equation and balance the formula equation giving the same number of each atom on both sides of the equation.

Balance the reaction between magnesium and oxygen given below:

$$Mg(s) + O_2(g) \rightarrow MgO(s)$$

â Determine the two changes in oxidation number.

$$Mg \rightarrow Mg^{2+}$$

$$O_2 \rightarrow O^{2-}$$

ã Balance the two half reactions for both mass and charge. Which is it oxidation and which is reduction?

$$Mg \rightarrow Mg^{2+} + 2e^- \qquad \text{oxidation}$$

$$O_2 + 4e^- \rightarrow 2O^{2-} \qquad \text{reduction}$$

ä Using the lowest common denominator, multiply each half-reaction by the quantity which will make the number of electrons lost equal to the number of electrons gained.

$2(Mg \rightarrow Mg^{2+} + 2e^-)$	oxidation
$O_2 + 4e^- \rightarrow 2O^{2-}$	reduction

å Add the two half-reactions together to obtain the net ionic equation.

$2Mg \rightarrow 2Mg^{2+} + 4e^-$	oxidation
$O_2 + 4e^- \rightarrow 2O^{2-}$	reduction

$$2Mg + O_2 \rightarrow 2Mg^{2+} + 2O^{2-}$$

æ Transfer the coefficients from the net ionic equation to the original formula equation and balance the original formula equation.

$$Mg(s) + O_2(g) \rightarrow MgO(s)$$

$$2Mg + O_2 \rightarrow 2MgO$$

Balance the following equation representing the reaction between copper and sulfuric acid.

$$Cu + H_2SO_4 \; 6 \; CuSO_4 + SO_2 + H_2O$$

â Determine the two changes in oxidation number.

$$Cu^0 \; 6 \; Cu^{2+}$$

$$S^{6+} \; 6 \; S^{4+}$$

ã Balance the two half reactions for both mass and charge. Which is it oxidation and which is reduction?

$$Cu^0 \; 6 \; Cu^{2+} + 2e^- \qquad \text{oxidation}$$

$$S^{6+} + 2e^- \; 6 \; S^{4+} \qquad \text{reduction}$$

ä Using the lowest common denominator, multiply each half-reaction by the quantity which will make the number of electrons lost equal to the number of electrons gained.

$$1(Cu^0 \rightarrow Cu^{2+} + 2e^-)$$
$$1(S^{6+} + 2e^- \rightarrow S^{4+})$$

å Add the two half-reactions together to obtain the net ionic equation.

$$Cu^0 \rightarrow Cu^{2+} + 2e^-$$
$$S^{6+} + 2e^- \rightarrow S^{4+}$$

$$Cu^0 + S^{6+} \rightarrow Cu^{2+} + S^{4+}$$

æ Transfer the coefficients from the net ionic equation to the original formula equation and balance the original formula equation.

$$Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$$
$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

Balance the following redox equation:

$$NH_3 + NaBrO \text{ 6 } NaBr + H_2O + N_2$$

â Determine the two changes in oxidation number.

$$N^{3-} \text{ 6 } N_2^{\,0}$$

$$Br^{1+} \text{ 6 } Br^{1-}$$

ã Balance the two half reactions for both mass and charge. Which is it oxidation and which is reduction?

$2N^{3-} \text{ 6 } N_2^{\,0} + 6e^-$	oxidation
$Br^{1+} + 2e^- \text{ 6 } Br^{1-}$	reduction

ä Using the lowest common denominator, multiply each half-reaction by the quantity which will make the number of electrons lost equal to the number of electrons gained.

$$2N^{3-} \rightarrow N_2^{0} + 6e^{-} \qquad \text{oxidation}$$
$$3(Br^{1+} + 2e^{-} \rightarrow Br^{1-}) \qquad \text{reduction}$$

å Add the two half-reactions together to obtain the net ionic equation.

$$2N^{3-} \rightarrow N_2^{0} + 6e^{-}$$
$$3Br^{1+} + 6e^{-} \rightarrow 3Br^{1-}$$

$$2N^{3-} + 3Br^{1+} \rightarrow N_2 + 3Br^{1-}$$

æ Transfer the coefficients from the net ionic equation to the original formula equation and balance the original formula equation.

$$NH_3 + NaBrO \rightarrow NaBr + H_2O + N_2$$
$$2NH_3 + 3NaBrO \rightarrow 3NaBr + 3H_2O + N_2$$

Balance the following redox equation:

$$FeSO_4 + H_2SO_4 + K_2Cr_2O_7 \text{ 6}$$

$$Fe_2(SO_4)_3 + K_2SO_4 + Cr_2(SO_4)_3 + H_2O$$

â Determine the two changes in oxidation number.

$$Fe^{2+} \text{ 6 } Fe^{3+}$$

$$Cr^{6+} \text{ 6 } Cr^{3+}$$

ã Balance the two half reactions for both mass and charge. Which is it oxidation and which is reduction?

$$2Fe^{2+} \text{ 6 } 2Fe^{3+} + 2e^- \qquad \text{oxidation}$$

$$2Cr^{6+} + 6e^- \text{ 6 } 2Cr^{3+} \qquad \text{reduction}$$

ä Using the lowest common denominator, multiply each half-reaction by the quantity which will make the number of electrons lost equal to the number of electrons gained.

$$3(2Fe^{2+} \; 6 \; 2Fe^{3+} + 2e^{-})$$
$$2Cr^{6+} + 6e^{-} \; 6 \; 2Cr^{3+}$$

å Add the two half-reactions together to obtain the net ionic equation.

$$6Fe^{2+} \; 6 \; 6Fe^{3+} + 6e^{-}$$
$$2Cr^{6+} + 6e^{-} \; 6 \; 2Cr^{3+}$$

$$6Fe^{2+} + 2Cr^{6+} \; 6 \; 6Fe^{3+} + 2Cr^{3+}$$

æ Transfer the coefficients from the net ionic equation to the original formula equation and balance the original formula equation.

$$FeSO_4 + H_2SO_4 + K_2Cr_2O_7 \; 6$$
$$Fe_2(SO_4)_3 + K_2SO_4 + Cr_2(SO_4)_3 + H_2O$$
$$6FeSO_4 + 7H_2SO_4 + K_2Cr_2O_7 \; 6$$
$$3Fe_2(SO_4)_3 + K_2SO_4 + Cr_2(SO_4)_3 + 7H_2O$$

Balance the following redox equation:

$$MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$$

â Determine the two changes in oxidation number.

$$Mn^{4+} \rightarrow Mn^{2+}$$

$$Cl^{1-} \rightarrow Cl_2^{0}$$

ã Balance the two half reactions for both mass and charge. Which is it oxidation and which is reduction?

$$Mn^{4+} + 2e^{-} \rightarrow Mn^{2+} \qquad \text{reduction}$$

$$2Cl^{1-} \rightarrow Cl_2^{0} + 2e^{-} \qquad \text{oxidation}$$

ä Using the lowest common denominator, multiply each half-reaction by the quantity which will make the number of electrons lost equal to the number of electrons gained.

$$Mn^{4+} + 2e^{-}\ 6\ Mn^{2+}$$
$$2Cl^{1-}\ 6\ Cl_2^{0} + 2e^{-}$$

å Add the two half-reactions together to obtain the net ionic equation.

$$Mn^{4+} + 2e^{-}\ 6\ Mn^{2+}$$
$$2Cl^{1-}\ 6\ Cl_2^{0} + 2e^{-}$$

$$Mn^{4+} + 2Cl^{1-}\ 6\ Cl_2 + Mn^{2+}$$

æ Transfer the coefficients from the net ionic equation to the original formula equation and balance the original formula equation.

$$MnO_2 + HCl\ 6\ MnCl_2 + H_2O + Cl_2$$

$$MnO_2 + 4HCl\ 6\ MnCl_2 + 2H_2O + Cl_2$$

Balance the following redox equation:

$As_2O_3 + HBrO_3 + H_2O$ 6 $H_3AsO_4 + HBr$

â Determine the two changes in oxidation number.

As^{3+} 6 As^{5+}

Br^{5+} 6 Br^{1-}

ã Balance the two half reactions for both mass and charge. Which is it oxidation and which is reduction?

$2As^{3+}$ 6 $2As^{5+} + 4e^-$	oxidation
$Br^{5+} + 6e^-$ 6 Br^{1-}	reduction

ä Using the lowest common denominator, multiply each half-reaction by the quantity which will make the number of electrons lost equal to the number of electrons gained.

$$3(2As^{3+} \rightarrow 2As^{5+} + 4e^{-})$$
$$2(Br^{5+} + 6e^{-} \rightarrow Br^{1-})$$

å Add the two half-reactions together to obtain the net ionic equation.

$$6As^{3+} \rightarrow 6As^{5+} + 12e^{-}$$
$$2Br^{5+} + 12e^{-} \rightarrow 2Br^{1-}$$

$$6As^{3+} + 2Br^{5+} \rightarrow 6As^{5+} + 2Br^{1-}$$

æ Transfer the coefficients from the net ionic equation to the original formula equation and balance the original formula equation.

$$As_2O_3 + HBrO_3 + H_2O \rightarrow H_3AsO_4 + HBr$$

$$3As_2O_3 + 2HBrO_3 + 9H_2O \rightarrow 6H_3AsO_4 + 2HBr$$

Oxidizing agent - The oxidizing agent is the species reduced in an oxidation-reduction reaction.

Oxidizing agents promote oxidation by gaining electrons.

Nonmetals generally make good oxidizing agents.

The best oxidizing agents are found on the top left of Chart N.

Reducing agent - The reducing agent is the species oxidized in an oxidation-reduction reaction.

Reducing agents promote reduction by losing electrons.

Metals generally make good reducing agents. The best reducing agents are found on the bottom right of Chart N.

Given the unbalanced equation:

$Cu + HNO_3$ 6 $Cu(NO_3)_2 + NO + H_2O$

Determine each of the following:

â The oxidation half-reaction balanced in terms of both atoms and electrons.

ã The reduction half reaction balanced in terms of both atoms and electrons.

ä The net ionic equation for the reaction.

å The complete, balanced chemical equation.

æ The species oxidized and reduced.

ç The oxidizing agent and the reducing agent.

$Cu + HNO_3$ 6 $Cu(NO_3)_2 + NO + H_2O$

â Cu^0 6 $Cu^{2+} + 2e^-$ oxidation

ã $N^{5+} + 3e^-$ 6 N^{2+} reduction

ä $3(Cu^0$ 6 $Cu^{2+} + 2e^-)$
$2(N^{5+} + 3e^-$ 6 $N^{2+})$

$3Cu^0$ 6 $3Cu^{2+} + 6e^-$
$2N^{5+} + 6e^-$ 6 $2N^{2+}$

$3Cu^0 + 2N^{5+}$ 6 $3Cu^{2+} + 2N^{2+}$

å $3Cu + 8HNO_3$ 6$3Cu(NO_3)_2 + 2NO + 4H_2O$

æ Cu^0 in oxidized, ˆ Cu^0 is the reducing agent

ç N^{5+} is reduced, ˆ HNO_3 is the oxidizing agent

Given the unbalanced equation:

$$KMnO_4 + HCl \ 6 \ KCl + MnCl_2 + H_2O + Cl_2$$

Determine each of the following:

â The oxidation half-reaction balanced in terms of both atoms and electrons.

ã The reduction half reaction balanced in terms of both atoms and electrons.

ä The net ionic equation for the reaction.

å The complete, balanced chemical equation.

æ The species oxidized and reduced.

ç The oxidizing agent and the reducing agent.

$$KMnO_4 + HCl \; 6 \; KCl + MnCl_2 + H_2O + Cl_2$$

â $Mn^{7+} + 5e^- \; 6 \; Mn^{2+}$ reduction

ã $2Cl^{1-} \; 6 \; Cl_2{}^0 + 2e^-$ oxidation

ä $2(Mn^{7+} + 5e^- \; 6 \; Mn^{2+})$

$5(2Cl^{1-} \; 6 \; Cl_2{}^0 + 2e^{-)}$

$2Mn^{7+} + 10e^- \; 6 \; 2Mn^{2+}$

$10Cl^{1-} \; 6 \; 5Cl_2{}^0 + 10e^-$

$2Mn^{7+} + 10Cl^{1-} \; 6 \; 2Mn^{2+} + 5Cl_2{}^0$

å $2KMnO_4 + 16HCl \; 6$

$2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

æ Cl^- is oxidized, ˆ HCl is the reducing agent.

ç Mn^{7+} is reduced, ˆ $KMnO_4$ is the oxidizing agent.

Given the unbalanced equation:

$MnSO_4 + KOH + O_2$ 6 $Mn_2O_3 + K_2SO_4 + H_2O$

Determine each of the following:

â The oxidation half-reaction balanced in terms of both atoms and electrons.

ã The reduction half reaction balanced in terms of both atoms and electrons.

ä The net ionic equation for the reaction.

å The complete, balanced chemical equation.

æ The species oxidized and reduced.

ç The oxidizing agent and the reducing agent.

$MnSO_4 + KOH + O_2$ 6 $Mn_2O_3 + K_2SO_4 + H_2O$

â $2Mn^{2+}$ 6 $2Mn^{3+} + 2e^-$ oxidation

ã $O_2^0 + 4e^-$ 6 $2O^{2-}$ reduction

ä $2(2Mn^{2+}$ 6 $2Mn^{3+} + 2e^{-)}$

$O_2^0 + 4e^-$ 6 $2O^{2-}$

$4Mn^{2+}$ 6 $4Mn^{3+} + 4e^-$

$O_2^0 + 4e^-$ 6 $2O^{2-}$

$4Mn^{2+} + O_2^0$ 6 $4Mn^{3+} + 2O^{2-}$

å $4MnSO_4 + 8KOH + O_2$ 6

$2Mn_2O_3 + 4K_2SO_4 + 4H_2O$

æ Mn^{2+} is oxidized, ˆ $MnSO_4$ is the reducing agent

ç O_2 is reduced, ˆ O_2 is the oxidizing agent

Given the unbalanced equation:

$$S + H_2O + Cl_2 \text{ 6 } H_2SO_4 + HCl$$

Determine each of the following:

â The oxidation half-reaction balanced in terms of both atoms and electrons.

ã The reduction half reaction balanced in terms of both atoms and electrons.

ä The net ionic equation for the reaction.

å The complete, balanced chemical equation.

æ The species oxidized and reduced.

ç The oxidizing agent and the reducing agent.

$$S + H_2O + Cl_2 \text{ 6 } H_2SO_4 + HCl$$

â $S^0 \text{ 6 } S^{6+} + 6e^-$ oxidation

ã $Cl_2{}^0 + 2e^- \text{ 6 } 2Cl^{1-}$ reduction

ä $S^0 \text{ 6 } S^{6+} + 6e^-$

$3(Cl_2{}^0 + 2e^- \text{ 6 } 2Cl^{1-})$

$$S^0 \text{ 6 } S^{6+} + 6e^-$$

$$3Cl_2{}^0 + 6e^- \text{ 6 } 6Cl^{1-}$$

$$3Cl_2{}^0 + S^0 \text{ 6 } S^{6+} + 6Cl^{1-}$$

å $S + 4H_2O + 3Cl_2 \text{ 6 } H_2SO_4 + 6HCl$

æ S^0 in oxidized, ˆ S^0 is the reducing agent

ç $Cl_2{}^0$ is reduced, ˆ Cl_2 is the oxidizing agent

Given the unbalanced equation:

$H_2SO_4 + K_2Cr_2O_7 + KI$ 6
$Cr_2(SO_4)_3 + K_2SO_4 + H_2O + I_2$

Determine each of the following:

â The oxidation half-reaction balanced in terms of both atoms and electrons.

ã The reduction half reaction balanced in terms of both atoms and electrons.

ä The net ionic equation for the reaction.

å The complete, balanced chemical equation.

æ The species oxidized and reduced.

ç The oxidizing agent and the reducing agent.

$H_2SO_4 + K_2Cr_2O_7 + KI$ 6

$Cr_2(SO_4)_3 + K_2SO_4 + H_2O + I_2$

â $2Cr^{6+} + 6e^-$ 6 $2Cr^{3+}$ reduction

ã $2I^-$ 6 $I_2^0 + 2e^-$ oxidation

ä $2Cr^{6+} + 6e^-$ 6 $2Cr^{3+}$

$3(2I^-$ 6 $I_2^0 + 2e^-)$

$2Cr^{6+} + 6e^-$ 6 $2Cr^{3+}$

$6I^-$ 6 $3I_2^0 + 6e^-$

$2Cr^{6+} + 6I^-$ 6 $2Cr^{3+} + 3I_2^0$

å $7H_2SO_4 + K_2Cr_2O_7 + 6KI$ 6

$Cr_2(SO_4)_3 + 4K_2SO_4 + 7H_2O + 3I_2$

æ Cr^{6+} is reduced, ˆ $K_2Cr_2O_7$ is the oxidizing agent

ç I^- in oxidized, ˆ KI is the reducing agent

Given the unbalanced equation:

$Al + HNO_3$ 6 $Al(NO_3)_3 + NH_4NO_3 + H_2O$

Determine each of the following:

â The oxidation half-reaction balanced in terms of both atoms and electrons.

ã The reduction half reaction balanced in terms of both atoms and electrons.

ä The net ionic equation for the reaction.

å The complete, balanced chemical equation.

æ The species oxidized and reduced.

ç The oxidizing agent and the reducing agent.

$Al + HNO_3$ 6 $Al(NO_3)_3 + NH_4NO_3 + H_2O$

â Al^0 6 $Al^{3+} + 3e^-$ oxidation

ã $N^{5+} + 8e^-$ 6 N^{3-} reduction

ä $8(Al^0$ 6 $Al^{3+} + 3e^-)$
$3(N^{5+} + 8e^-$ 6 $N^{3-)}$

$8Al^0$ 6 $8Al^{3+} + 24e^-$
$3N^{5+} + 24e^-$ 6 $3N^{3-}$

$8Al^0 + 3N^{5+}$ 6 $8Al^{3+} + 3N^{3-}$

å $8Al + 30HNO_3$ 6
$8Al(NO_3)_3 + 3NH_4NO_3 + 9H_2O$

æ Al^0 is oxidized, ˆ Al is the reducing agent

ç N^{5+} in reduced, ˆ HNO_3 is the oxidizing agent

CHART N

â All half reactions on Chart N are written as reduction half-reactions.

ã The species on the top left of Chart N are the best oxidizing agents. Fluorine in the best oxidizing agent on Chart N.

ä The species on the bottom right of Chart N are the best reducing agents. Lithium is the best reducing agent on Chart N.

å When using Chart N to predict spontaneous chemical reactions, the top half-reaction is always the reduction half- reaction and the bottom half-reaction is always the oxidation half-reaction.