SCH4U7

Redox

Introduction to Redox Reactions

OXIDATION

- Historically: to combine with oxygen
  i.e. iron + oxygen → rust (iron oxide)
- To increase the bonds to oxygen

REDUCTION

- Historically: formation of a metal from its compounds i.e. iron ore → iron
- To decrease the bonds to oxygen

Current definitions are based on <u>electron transfers</u>:

LOSS of electrons is OXIDATION, GAIN of electrons is REDUCTION

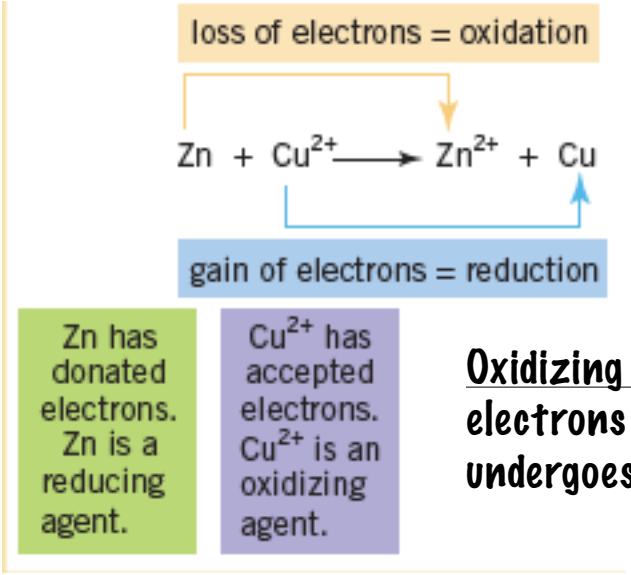
OIL RIG **LEO the lion says GER**

Redox reactions MUST occur together, one substance loses electrons, another gains.

<u>Example:</u> Consider the single displacement reaction: $Zn_{(s)} + CuSO_{4(aq)} \dashrightarrow Cu_{(s)} + ZnSO_{4(aq)}$

Write the total ionic equation: $Zn_{(s)} + Cu^{2+}_{(aq)} + SO_4{}^{2-}_{(aq)} \dashrightarrow Zn^{2+}_{(aq)} + Cu_{(s)} + SO_4{}^{2-}_{(aq)}$

Excluding the spectator ion:



<u>Reducing Agent:</u> donates electrons in a redox rxn and undergoes oxidation

<u>Oxidizing Agent:</u> accepts electrons in a redox rxn and undergoes reduction

This example was for an ionic reaction, where there is a clear exchange of electrons.

What about $C(s) + O_2(g) \rightarrow CO_2(g)$

Carbon is oxidized (since it has more bonds to oxygen) but how were the electrons transferred?

We use oxidation numbers to deal with these situations. They are not real charges.

RULES FOR ASSIGNING OXIDATION NUMBERS

TABLE 10.1.1 ASSIGNING OXIDATION NUMBERS

Rule for assigning oxidation numbers	Examples of application of the rule
1 The oxidation number of each atom in a pure element is zero.	Zn, O in O_2, and P in P_4 all have an oxidation number of zero.
2 The oxidation number of an atom in a monatomic ion is equal to the charge on the ion.	Na^+ has an oxidation number of +1. S^{2-} has an oxidation number of –2.
3 In compounds containing oxygen, each oxygen atom has an oxidation number of –2 (exceptions are OF_2, in which the oxidation number of oxygen is +2, and peroxides (O_2^{2-}), in which it is –1).	In H_2O and CO_2 each oxygen atom has an oxidation number of –2. In H_2O_2 each oxygen atom has an oxidation number of –1.
4 In compounds containing hydrogen, each hydrogen atom has an oxidation number of +1 (exceptions are metal hydrides such as NaH, in which the oxidation number of hydrogen is –1).	In NH_3 and H_2O each hydrogen atom has an oxidation number of +1. In KH the hydrogen atom has an oxidation number of –1.
5 For a molecule, the sum of the oxidation numbers of the atoms equals zero.	The sum of the oxidation numbers of the atoms in CH_4 is zero. As each hydrogen atom has an oxidation number of +1, the oxidation number of the carbon atom is –4: $(x + (4 \times +1) = 0, x = -4)$
6 For a polyatomic ion, the sum of the oxidation numbers of the atoms equals the charge on the ion.	The sum of the oxidation numbers of the atoms in PO_4^{3-} is –3. As each oxygen atom has an oxidation number of –2, the oxidation number of the phosphorus atom is +5: $(x + (4 \times -2) = -3, x = +5)$
7 In a compound, the most electronegative atom is assigned the negative oxidation number.	In SF_6, the oxidation number of each fluorine atom is –1. The oxidation number of the sulfur atom is +6: $(x + (6 \times -1) = 0, x = +6)$

Determine the oxidation number of each element in each of the following species:

a H_2S **b** H_2SO_4 **c** $KClO_4$

Determine the oxidation number of each element in each of the following species.

a MgO **b** H_2O **c** HPO_4^- **d** C_2H_4

e N_2O_4 **f** $KMnO_4$

Determine the oxidation number of the metal in each of the following metal oxides and hence write the name of the compound using Roman numerals to show the oxidation number of the metal.

a CuO **b** Cu_2O **c** MnO_2 **d** V_2O_5

We can use oxidation numbers to determine which reactant is being oxidized or reduced.

Oxidation is an increase in oxidation number.

Reduction is a decrease in oxidation number.

Use oxidation numbers to determine which reactant has been oxidized and which has been reduced.

$$SO_2(g) + NO_2(g) \rightarrow SO_3(g) + NO(g)$$

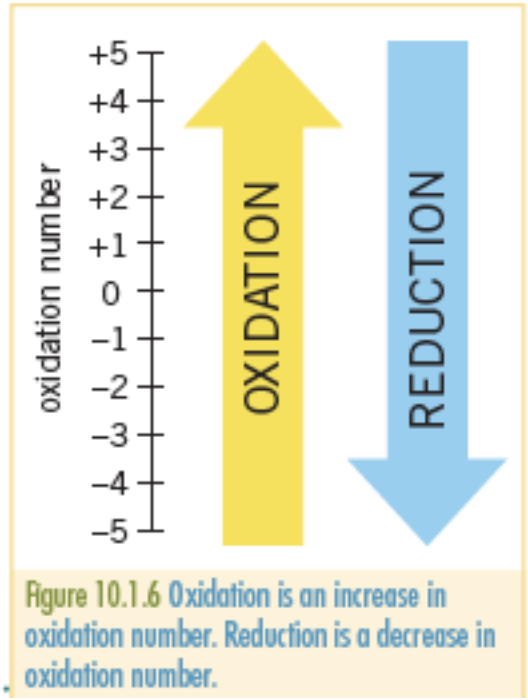


Figure 10.1.6 Oxidation is an increase in oxidation number. Reduction is a decrease in oxidation number.

For each of the questions given below, use oxidation numbers to determine whether the bolded chemical is being oxidized (O), reduced (R), neither oxidized nor reduced (N), or both oxidized and reduced (B).

a $\mathbf{Cl_2}(g) + 2e^- \rightarrow 2Cl^-(aq)$

b $\mathbf{Ni}(s) \rightarrow Ni^{2+}(aq) + 2e^-$

c $\mathbf{Ba^{2+}}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

d $2\mathbf{Fe^{2+}}(aq) \rightarrow Fe(s) + Fe^{3+}(aq)$

Use oxidation numbers to show that the respiration reaction shown in the equation below is a redox reaction.

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

Use oxidation numbers to identify which reactant has been oxidized in each of the following redox reactions.

a $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

b $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \rightarrow 2Cr^{3+}(aq) + 3I_2(l) + 7H_2O(l)$

c $MnO_2(s) + SO_2(aq) \rightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$

SCH4U7	# Balancing Redox Equations
Redox	

NO shortcuts! This format must be followed in the given sequence.

Step 1 Write separate equations for the oxidation and reduction half-reactions.

Step 2 For each half reaction:

a) Balance all the elements except hydrogen and oxygen
b) Balance oxygen using H_2O.
c) Balance hydrogen using H^+ .
d) Balance the charge using electrons.

Step 3 If necessary, multiply one or both balanced half-reactions by an integer to equalize the number of electrons transferred in the two half-reactions. # electrons lost = # electrons gained.

Step 4 Add the half-reactions, and cancel identical species.

Step 5 Check that the elements and charges are balanced.

If the redox reaction is occurring in **a basic solution** continue on.

Step 6 To both sides of the equation obtained above, add a number of OH^- ions that is equal to the number of H^+ ions. (one wants to eliminate H^+ by forming water.)

Step 7 Form H_2O on the side containing H^+ and OH^- ions, and eliminate the number of H_2O molecules that appear on both sides of the equation.

Step 8 Check that the elements and charges are balanced.

Let's try an example:

In acidic solution, balance $\mathbf{ClO_4^- + NO_2 \dashrightarrow Cl^- + NO_3^-}$

Assign oxidation numbers.

Is it a redox?

Which was Oxidized?

Which was Reduced?

Write the half reactions separately.

OXIDIZED **REDUCED**

Balance other atoms

Balance with H_2O

Balance with H^+

Balance with e^-

$$8NO_2 + ClO_4^- + H_2O \dashrightarrow 8NO_3^- + 8H^+ + Cl^-$$

In basic conditions, balance $Br_2 \dashrightarrow BrO_3^- + Br^-$

Assign oxidation numbers.

Is it a redox?

Which was Oxidized?

Which was Reduced?

Write the half reactions separately.

OXIDIZED **REDUCED**

Balance other atoms

Balance with H_2O

Balance with H^+

Balance with e^-

REDOX WORKSHEET

Balance the following redox reactions. Assume all are in acidic solutions unless otherwise indicated.

1. $NO_3^- + Cu \rightarrow NO_2 + Cu^+$
2. $IO_3^- + AsO_3^{3-} \rightarrow I^- + AsO_4^{3-}$
3. $SO_4^{2-} + Zn \rightarrow Zn^{2+} + SO_2$
4. $NO_3^{1-} + Zn \rightarrow NH_4^+ + Zn^{2+}$
5. $Cr^{3+} + BiO_3^{1-} \rightarrow Cr_2O_7^{2-} + Bi^{3+}$
6. $I_2 + OCl^{1-} \rightarrow IO_3^{1-} + Cl^{1-}$
7. $Mn^{2+} + BiO_3^{1-} \rightarrow MnO_4^{1-} + Bi^{3+}$
8. $MnO_4^{1-} + C_2O_4^{2-} \rightarrow CO_2 + MnO_2$ (basic)
9. $ClO_3^{1-} + N_2H_4 \rightarrow NO + Cl^{1-}$
10. $NiO_2 + Mn(OH)_2 \rightarrow Mn_2O_3 + Ni(OH)_2$
11. $SO_3^{2-} + CrO_4^{2-} \rightarrow SO_4^{2-} + CrO_2^{1-}$ (basic)
12. $Au + CN^{1-} + O_2 \rightarrow Au(CN)_4^{1-} + OH^{1-}$ (basic)