Oxidation Reduction Reactions

Model 1 – Assigning Oxidation Numbers

Oxidation Number (O.N.): The charge or apparent charge that an atom in a molecular compound or ion would have if all of the electrons in its bonds belonged entirely to the more electronegative atom—i.e. the atom that more strongly attracts the shared electrons.

Guidelines for Assigning Oxidation Numbers

1) In **free elements** (i.e., in the uncombined state), each atom has an oxidation number of zero—e.g. Na, Zn, N_2, S_8, P_4, F_2

2) The oxidation number of **monatomic ions** is equal to the charge on the ion—e.g. the O.N. of Na^+ is +1, the O.N. of Cl^- ion is -1, that of Mg^{2+} is +2.

3) The oxidation number of **fluorine** in all of its compounds is -1. Fluorine is the most electronegative element.

4) The oxidation number of **oxygen** in most compounds (called "oxides") is **-2**—e.g. Na_2O, ZnO, SO_2, SO_3. In peroxides (e.g. Na_2O_2, H_2O_2) the oxidation number of oxygen is -1. In superoxides (O_2^-), the oxidation number of oxygen is - 1 /2. (When oxygen is combined with fluorine it will have a positive oxidation number!)

5) The oxidation number of **hydrogen is +1**, except when it is bonded to metals in binary compounds called hydrides in which its oxidation number is -1—e.g. NaH, CaH_2, AlH_3.

6) The **sum of the oxidation numbers in a neutral compound is zero**: H_2O: 2(+1) + (-2) = 0

7) The **sum of the oxidation numbers in a polyatomic ion is equal to the charge on the ion**. The oxidation number of the sulfur atom in the SO_4^{2-} ion must be +6, for example, because the sum of the oxidation numbers of the atoms in this ion must equal -2: SO_4^{2-} : (+6) + 4(-2) = -2

KEY QUESTIONS

1. What is the oxidation number of a pure element such as N_2, O_2, Ne, S_8, P_4, Hg, Na or Fe?

2. For the polyatomic ion $Cr_2O_7^{2-}$, what is the sum of the oxidation numbers of chromium and oxygen atoms?

3. In terms of relative electronegativities, why is the oxidation number of hydrogen +1 in compounds with C, N, O, and Cl, but –1 with metals like Na, K, Mg, and Al?

4. In terms of relative electronegativities, why does chlorine have a negative oxidation umber in CCl_4, but a positive oxidation number in $HClO_4$?

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EXERCISES

Using the guidelines above and the insight gained from your answers to the Key Questions, assign oxidation numbers to all atoms in the following compounds or ions:

Li_2O	P_4	CaH_2	Fe_2O_3	$MnCl_3$
_____	_____	_____	_____	_____
O_2	H_2O	H_2O_2	OF_2	S_8
_____	_____	_____	_____	_____
$HClO_2$	Cs_2O_2	CH_2Cl_2	ClO_4^-	PO_3^{3-}
_____	_____	_____	_____	_____
MnO_4^-	MnO_4^{2-}	$(NH_4)_3PO_4$	Na_2SO_3	$Pb(IO_4)_2$
_____	_____	_____	_____	_____

PRACTICE*!*

There are numerous compounds and polyatomic ions containing nitrogen and oxygen. None of these are peroxides or superoxides. Complete the following table.

Species	Oxidation number of each oxygen	Total charge of all oxygen atoms	Total charge of all nitrogen atoms	Average oxidation number of each nitrogen atom
NO				
NO_2				
N_2O				
NO_2^-				
NO_3^-				

Oxidation Reduction Reactions

Read This*!*

The process of oxidation and reduction can be thought of as a transfer of electrons from one atom to another. Thus, one atom gives up electrons and the other atom gains them. As a result of this process, the oxidation numbers of both atoms change. All redox reactions can be divided up into two reactions—an **oxidation half-reaction** and a **reduction half-reaction**. This allows for better understanding of the electron transfer process.

Model 2 – Half Reactions

A. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

ox: $Zn \rightarrow Zn^{2+} + 2e^-$

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

B. $2I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_2(s) + 2SO_4^{2-}(aq)$

ox: $2I^- \rightarrow I_2 + 2e^-$

red: $S_2O_8^{2-} + 2e^- \rightarrow 2SO_4^{2-}$

C. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

ox: $Fe \rightarrow Fe^{3+} + 3e^-$

red: $O_2 + 4e^- \rightarrow 2O^{2-}$

1. What does the "e–" symbol represent in the oxidation and reduction half-reactions shown in Model 2?

2. Look at the oxidation half-reactions in Model 2.

a. Are electrons lost or gained by an atom during the process of oxidation?

b. Does the oxidation number of an atom involved in the process of oxidation increase or decrease?

3. Look at the examples of reduction in Model 2.

a. Are electrons lost or gained by an atom during the process of reduction?

b. Does the oxidation number of an atom involved in the process of reduction increase or decrease?

Oxidation Reduction Reactions

Read This!

Oxidation occurs when atoms lose electrons. **Reduction** occurs when atoms gain electrons. These two processes always occur together. In other words, you can't just let electrons loose into space—they must be grabbed by some other atom. Likewise, you can't just grab electrons from space—they must be taken from some other atom. An easy way to remember these processes is to remember the phrase "LEO the lion goes GER."

LEO = Loss of Electrons is Oxidation GER = Gain of Electrons is Reduction

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4. When iron is exposed to oxygen, it forms rust as described by the following equation:

$$4Fe(s) + 3O_2\,(g) \rightarrow 2Fe_2O_3\,(s)$$

a. What are the oxidation numbers of the atoms in all species?

b. What element was oxidized in the reaction above? What element was reduced?

c. What is the **oxidation half-reaction** and the **reduction half-reaction**?

d. In redox reactions, the reactant containing the element being reduced is called the *oxidizing agent* because it is removing electrons from another atom. The reactant containing the element being oxidized is called the *reducing agent* because it is providing electrons to another atom.

   What is the reducing agent in the reaction above? What is the oxidizing agent?

e. How can you distinguish a redox reaction from other types of chemical reactions?

Oxidation Reduction Reactions

EXERCISES*!*

Assign oxidation numbers to all the atoms in the following reactions. For the redox reactions, identify the atom oxidized, the atom reduced, the oxidizing agent, and the reducing agent.

a)

$$Pb(NO_3)_2\ (aq) + 2NaI(aq) \rightarrow PbI_2\ (s) + 2NaNO_3\ (aq)$$

Redox (yes/no)	**Atom Oxidized**	**Atom Reduced**	**Oxidizing Agent**	**Reducing Agent**

b)

$$2H_2O \rightarrow 2H_2 + O_2$$

Redox (yes/no)	**Atom Oxidized**	**Atom Reduced**	**Oxidizing Agent**	**Reducing Agent**

c)

$$CH_4 + 2O_2 \rightarrow 2H_2O + CO_2$$

Redox (yes/no)	**Atom Oxidized**	**Atom Reduced**	**Oxidizing Agent**	**Reducing Agent**

d)

$$HCl + NaOH \rightarrow H_2O + NaCl$$

Redox (yes/no)	Atom Oxidized	Atom Reduced	Oxidizing Agent	Reducing Agent

e)

$$Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$$

Redox (yes/no)	Atom Oxidized	Atom Reduced	Oxidizing Agent	Reducing Agent

f)

$$6FeSO_4(aq) + K_2Cr_2O_7(aq) + 7H_2SO_4(aq) \rightarrow$$

$$Cr_2(SO_4)_3(aq) + 3Fe_2(SO_4)_3(aq) + K_2SO_4(aq) + 7H_2O(l)$$

Redox (yes/no)	Atom Oxidized	Atom Reduced	Oxidizing Agent	Reducing Agent