Predicting Redox Reactions

A redox reaction may be explained as a transfer of electrons from one substance to another. Since two particles must be involved in an electron transfer, this transfer can be thought of as a competition for electrons.

Using a tug-of-war analogy, each particle pulls on the same electrons. If one particle is able to pull electrons away from the other, a spontaneous reaction occurs. Otherwise, no reaction occurs.

For example, when copper metal is placed in a solution of silver nitrate, a blue solution of copper nitrate is produced along with metallic silver.

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

In this reaction, the silver ion (Ag^+) pulls harder on copper's electrons than copper does, thus winning the tug-of-war. Copper loses electrons, becoming oxidized to form copper ions (Cu^{2+}), while the silver ions gain electrons, becoming reduced to form metallic silver (Ag).

Unlike many other reactions that we have examined, this reaction is not reversible. If you place silver metal in a solution of copper nitrate, no reaction will occur.

$$Ag(s) + Cu(NO_3)_2(aq) \rightarrow no\ reaction$$

Because silver pulls harder on its electrons than the copper ions do, copper is unable to remove electrons from the silver atoms. Thus, no reaction occurs.

The above example described a single-replacement reaction. All reactions of this kind are redox reactions. Chemical reactions for a great many single-replacement reactions can be written, but not all of the reactions actually occur. The question that arises is, how can we predict whether or not a given reaction will occur spontaneously?

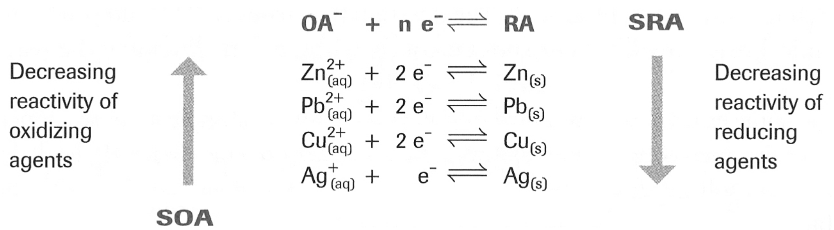
Redox Tables

A **redox table** consists of a series of balanced half-reactions, similar to the one shown below.

$$Ag^{+} + e^{-} \Leftrightarrow Ag(s)$$

Each half-reaction can be read as a reduction by reading it from left to right, or as an oxidation by reading it from right to left. Thus, the entity on the left (Ag^{+} in the above half-reaction) is the oxidizing agent, while the entity on the right (Ag) is the reducing agent.

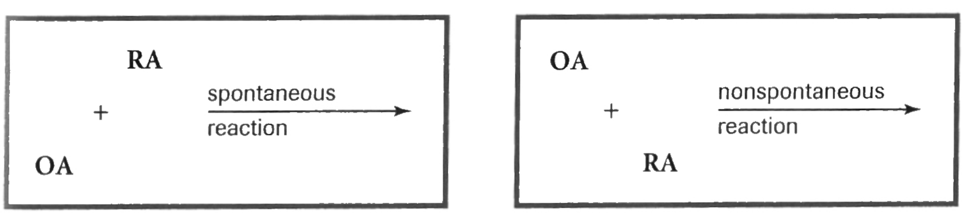
The half-reactions on a redox table are listed according to the relative strengths of the oxidizing and reducing agents, as illustrated by the sample table below.



On the left side, the oxidizing agents are listed in order from weakest (at the top) to strongest (at the bottom). On the right side, the reducing agents are listed in order from strongest (at the top) to weakest (at the bottom).

The Spontaneity Rule

The image below illustrates how you can use a redox table to predict whether or not a reaction is spontaneous.



A spontaneous redox reaction occurs only if the reducing agent (RA) is above the oxidizing agent (OA) on a table of relative strengths of oxidizing and reducing agents (redox table).

Example 1
Will the following reaction occur? If so, complete the reaction.

$$Cu(s) + Mg^{2+}(aq) \rightarrow$$

Example 2
Predict whether or not a reaction will occur when tin strips are placed in hydrochloric acid.

Redox tables are normally developed from experimental evidence. A series of reactions are attempted. For each attempt, it is recorded whether a reaction occurs or not. Once this evidence is collected, it can be used to construct a redox table. The following example illustrates this procedure.

Example 3
The following reactions were performed. Construct a redox table from this information.

$$3Co^{2+}(aq) + 2In(s) \rightarrow 2In^{3+}(aq) + 3Co(s)$$
$$Cu^{2+}(aq) + Co(s) \rightarrow Co^{2+}(aq) + Cu(s)$$
$$Cu^{2+}(aq) + Pd(s) \rightarrow no\ reaction$$

Worksheet

1. The following reactions were performed. Construct a redox table.

$$Co^{2+}(aq) + Zn(s) \rightarrow Co(s) + Zn^{2+}(aq)$$
$$Mg^{2+}(aq) + Zn(s) \rightarrow no\ reaction$$

2. In a school laboratory four metals were combined with each of four solutions. Construct a redox table.

$$Be(s) + Cd^{2+}(aq) \rightarrow Be^{2+}(aq) + Cd(s)$$
$$Cd(s) + 2H^{+}(aq) \rightarrow Cd^{2+}(aq) + H_2(g)$$
$$Ca^{2+}(aq) + Be(s) \rightarrow no\ reaction$$
$$Cu(s) + H^{+}(aq) \rightarrow no\ reaction$$

3. Use the reactions below to construct a redox table.

$$Ag(s) + Br_2(l) \rightarrow AgBr(s)$$
$$Ag(s) + I_2(s) \rightarrow no\ reaction$$
$$Cu^{2+}(aq) + I^{-}(aq) \rightarrow no\ reaction$$
$$Br_2(l) + Cl^{-}(aq) \rightarrow no\ reaction$$

4. Arrange the following metal ions in order of decreasing strength as oxidizing agents: lead(II) ions, silver ions, zinc ions, and copper(II) ions.

5. Use the spontaneity rule to predict whether the following mixtures will show evidence of a reaction. Do not write the equation for the reaction.

   a) nickel metal in a solution of silver ions

   b) zinc metal in a solution of aluminum ions

   c) an aqueous mixture of copper(II) ions and iodide ions

   d) chlorine gas bubbled into a bromide ion solution

   e) an aqueous mixture of copper(II) ions and tin(II) ions

   f) copper metal in nitric acid

6. Prepare a redox table using the following experimentally obtained data. (R = reaction, NR = no reaction)

	Al^{3+}	Tl^{+}	Ga^{3+}	In^{3+}
Al	NR	R	R	R
Tl	NR	NR	NR	NR
Ga	NR	R	NR	R
In	NR	R	NR	NR