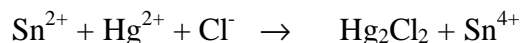


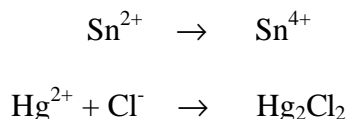
## BALANCING REDOX EQUATIONS BY THE ION-ELECTRON METHOD\*

In addition to the oxidation-number-change method discussed in the last section, there is still another method particularly well suited for balancing net ionic equations for oxidation-reduction reactions in solution. The procedure is called the **ion-electron method** and involves breaking the overall reaction into two **half-reactions**, one for the oxidation step and one for reduction. Each half-reaction is first balanced materially (that is, in terms of atoms), and then electrically by adding electrons to the side of the half-reaction deficient in negative charge. Finally, the balanced half-reactions are added together in such a way that the electrons cancel from both sides of the final equation.

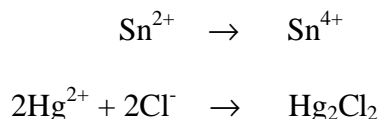
As a simple example, let's consider the reaction of  $\text{Sn}^{2+}$  with  $\text{Hg}^{2+}$  in the presence of chloride ion to produce  $\text{Hg}_2\text{Cl}_2$  and  $\text{Sn}^{4+}$  as products. The unbalanced equation is



*The first step is to divide the equation into the two half-reactions.* When you do this, it is important to remember that the atoms on each side of each half-reaction should be of the same kind. The half-reactions here are

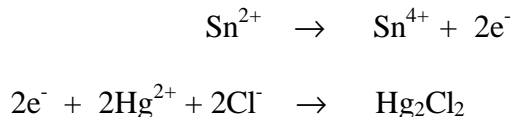


*Next, we balance each half-reaction in terms of atoms.* We don't have to do anything to the first one, but we must adjust the coefficients in the second.

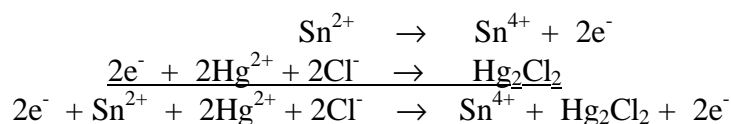


*The third step is to balance the charge by adding electrons to the more positive (or less negative) side.* In the first half-reaction the net charge on the left is 2+ and on the right it is 4+, so we add  $2e^-$  to the right so that the net charge on both sides is the same. In the second half-reaction, the net

charge is 2+ on the left [2 x (2+) + 2 (1-) = 2+] and zero on the right. Therefore, we add 2e<sup>-</sup> to the left. This gives



**Step four is to make the number of electrons gained equal the number lost.** In this example, we don't have to actually do anything because the necessary condition is already fulfilled. Now the half-reactions can be added together.



**Finally, we cancel anything that is the same on both sides.** In this and every other reaction balanced by this method the electrons must cancel. The finished equation is



Notice that both atoms and charge are in balance.

### ***Redox reactions in acidic solution***

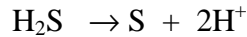
In many oxidation-reduction reactions that take place in aqueous solution, water plays an active role. Any aqueous solution contains the species H<sub>2</sub>O, H<sup>+</sup>, and OH<sup>-</sup>. In acidic solutions the predominant species are H<sub>2</sub>O and H<sup>+</sup>; in basic solutions they are H<sub>2</sub>O and OH<sup>-</sup>. When balancing half-reactions that occur in solution, we can use these species to achieve material balance. For example, let us consider the reaction between dichromate ion, Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>, and hydrogen sulfide in acidic solution to produce chromium(III) ion and elemental sulfur.



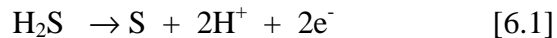
To balance the equation we write individual half-reactions for the changes that occur for chromium and sulfur. For the sulfur we write



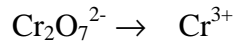
The half-reaction is balanced materially by adding two  $\text{H}^+$  ions to the right.



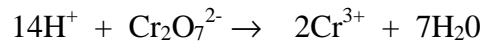
Electrical balance is achieved by adding two electrons to the right so that the net charge on both sides is zero.



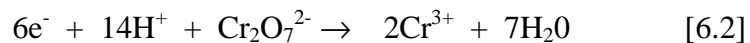
For chromium we begin by writing



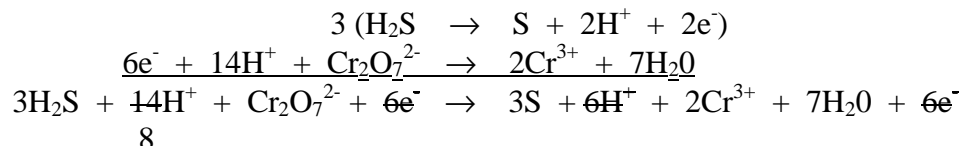
The chromium atoms are balanced first by placing the coefficient 2 before the  $\text{Cr}^{3+}$  on the right. The oxygens that appear on the left are balanced by adding  $\text{H}_2\text{O}$  (seven of them) to the right and the hydrogen imbalance that results is removed by placing  $14\text{H}^+$  on the left.



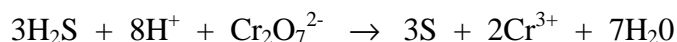
Now we have to balance the half-reaction electrically. The net charge on the left is  $14 \times (1+) + (2-) = 12+$ . On the right the net charge is  $6+$ . The algebraic difference between these is the number of electrons we add to the more positive side. Thus we need  $6\text{e}^-$  on the left.



The next step is to add these half-reactions (Equations 6.1 and 6.2) so that all the electrons in the final equation cancel. This is accomplished by multiplying Equation 6.1 through by 3 before adding.



The six electrons cancel, as do six of the  $\text{H}^+$ , so the final balanced equation reads



In summary, when balancing half-reactions in acid solution:

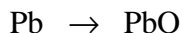
- (a) To balance a hydrogen atom we add a hydrogen ion,  $\text{H}^+$ , to the other side of the equation.
- (b) To balance an oxygen atom we add a water molecule to the side deficient in oxygen and then two  $\text{H}^+$  ions to the opposite side to remove the hydrogen imbalance.

### ***Redox reactions in basic solution***

We have just seen that  $\text{H}_2\text{O}$  and  $\text{H}^+$  are used to balance half-reactions that occur in acidic solution. In a basic solution the dominant species are  $\text{H}_2\text{O}$  and  $\text{OH}^-$ , so these are the species that should be used to achieve material balance.

Although you can use  $\text{H}_2\text{O}$  and  $\text{OH}^-$  directly (see footnote 3), the simplest technique is to first balance the reaction as if it occurred in acidic solution, and then perform the "conversion" described below to adjust it to conform to conditions in basic solution.

Suppose we wished to balance the following half-reaction taking place in a basic solution.



First we balance it as if it occurred in an acidic solution.

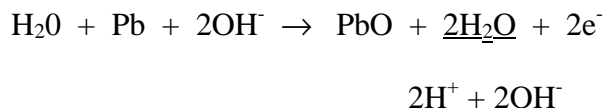


*The conversion to basic solution follows these three steps:*

**Step 1** For each  $H^+$  that must be eliminated from the equation, add an  $OH^-$  to both sides of the equation. In this example, we have to eliminate  $2H^+$ , so we add  $2OH^-$  to each side.



**Step 2** Combine  $H^+$  and  $OH^-$  to form  $H_2O$ . We have  $2H^+$  and  $2OH^-$  on the right, which give  $2H_2O$ .



**Step 3** Cancel any  $H_2O$  that are the same on both sides. We can cancel one  $H_2O$  from each side. The final balanced half-reaction in basic solution is



### **Footnotes**

\* Source: "General Chemistry: Principles and Structure, 3<sup>rd</sup> Ed." by James E. Brady and Gerard E. Humiston, John Wiley & Sons, New York 1982, pages 190-192

- 3 To balance half-reactions in basic solution, the following rules can be used:
- To balance a hydrogen atom, we add one  $H_2O$  molecule to the side of the half-reaction deficient in hydrogen, and to the other side we add one hydroxide ion.
  - To balance one oxygen atom, we add two hydroxide ions to the side deficient in oxygen, and one  $H_2O$  molecule to the other side.