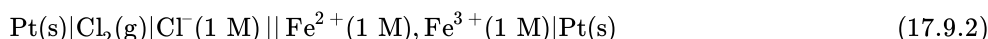## 17.9: Cell Notation and Conventions

Rather than drawing a complete diagram like the figures in the Galvanic Cells section, it is convenient to specify a galvanic cell in shorthand form. The two cells we have just described would be written as

$$\text{Zn}|\text{Zn}^{2+}(1 \text{ M})\,||\,\text{Cu}^{2+}(1 \text{ M})|\text{Cu} \tag{17.9.1}$$

and

$$\text{Pt(s)}|\text{Cl}_2(\text{g})|\text{Cl}^-(1 \text{ M})\,||\,\text{Fe}^{2+}(1 \text{ M}), \text{Fe}^{3+}(1 \text{ M})|\text{Pt(s)} \tag{17.9.2}$$

The components of the cell are written in order, starting with the left-hand and moving across the salt bridge to the right. A single vertical line indicates a phase boundary, such as that between the solid Zn electrode and $\text{Zn}^{2+}(\text{aq})$, or between $\text{Cl}_2(\text{g})$ and $\text{Cl}^-(\text{aq})$. The double vertical line represents a salt bridge. Spectator ions, like $\text{SO}_4^{2-}(\text{aq})$ in the Zn–Cu cell, are usually omitted.

By convention, the electrode written to the *left* of the salt bridge in this cell notation is always taken to be the *anode*, and the associated half-equation is always written as an *oxidation*. The right-hand electrode is therefore always the *cathode*, and the half-equation is always written as a *reduction*. This is easy to remember, because reading from left to right gives anode and cathode in alphabetical order. The **cell reaction** corresponding to a given shorthand description is obtained by summing the half-equations after multiplying by any factors necessary to equalize the number of electrons lost at the anode with the number gained at the cathode. Unless otherwise stated, standard conditions of 1 *M* are usually implied. However, non-standard cells can be created.
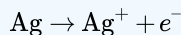
> ### ✔ Example 17.9.1: Half-Reactions
>
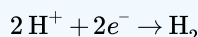> Write the half-equations and cell reactions for each of the following cells:
>
> a. $\text{Ag}|\text{Ag}^+\,||\,\text{H}^+\,|\,\text{H}_2\,|\,\text{Pt(s)}$
> b. $\text{Pt(s)}\,|\,\text{Cr}_2\text{O}_7^{2-}, \text{Cr}^{3+}, \text{H}^+\,||\,\text{Br}^-\,|\,\text{Br}_2(\text{l})\,|\,\text{Pt(s)}$
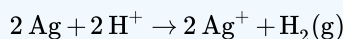>
> **Solution**
>
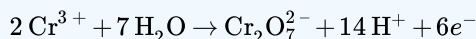> **a)** Since the half-equation at the left electrode is taken by convention to be oxidation, we have
>
> $$\text{Ag} \rightarrow \text{Ag}^+ + e^-$$
>
> At the right-hand electrode, then, we must have a reduction:
>
> $$2\,\text{H}^+ + 2e^- \rightarrow \text{H}_2$$
>
> Multiplying the first half-equation by 2 and summing gives the cell reaction
>
> $$2\,\text{Ag} + 2\,\text{H}^+ \rightarrow 2\,\text{Ag}^+ + \text{H}_2(\text{g})$$
>
> **b)** Following the same procedure as in part a, we obtain
>
> $$2\,\text{Cr}^{3+} + 7\,\text{H}_2\text{O} \rightarrow \text{Cr}_2\text{O}_7^{2-} + 14\,\text{H}^+ + 6e^-$$
>
> Left electrode
>
> $$(\text{Br}_2(\text{l}) + 2e^- \rightarrow 2\,\text{Br}^-) \times 3$$
>
> Right electrode
>
> $$2\,\text{Cr}^{3+} + 7\,\text{H}_2\text{O} + 3\,\text{Br}_2(\text{l}) \rightarrow \text{Cr}_2\text{O}_7^{2-} + 14\,\text{H}^+ + 6\,\text{Br}^-$$
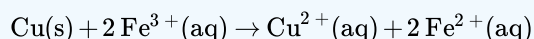>
> Cell reaction
>
> Note: The procedures described in other sections for balancing redox equations were used in arriving at the above results.
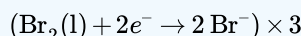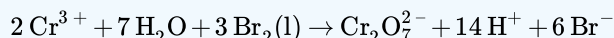
> ✔ **Example 17.9.2: Galvanic Cell**
>
> Describe in shorthand notation a galvanic cell for which the cell reaction is
>
> $$\text{Cu(s)} + 2\,\text{Fe}^{3+}(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\,\text{Fe}^{2+}(\text{aq})$$
>
> **Solution**
>
> First divide the cell reaction into half-equations:
>
> - Oxidation: $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2e^{-}$
> - Reduction: $\text{Fe}^{3+} + e^{-} \rightarrow \text{Fe}^{2+}$
>
> Then write the oxidation as the left-hand electrode and the reduction on the right:
>
> $$\text{Cu}|\text{Cu}^{2+} \parallel \text{Fe}^{2+}, \text{Fe}^{3+}|\text{Pt}$$
>
> (Since both Fe^{2+} and Fe^{3+} are in solution, a Pt electrode is used.)

The conventions we have developed can be used to decide whether the cell reaction is actually spontaneous. If it is, an oxidation will release electrons to the external circuit at the left-hand electrode. If a voltmeter is placed in the circuit, these electrons will make its left-hand terminal negative. Since the right-hand electrode corresponds to reduction, electrons will be withdrawn from the right-hand terminal of the voltmeter. This is shown for the Zn–Cu cell in Figure 1 from the Galvanic Cells section. You can readily confirm that the spontaneous cell reaction (Eq. (1) from Galvanic Cells) corresponds to the shorthand cell notation of Equation 17.9.1.

For the cell shown in Figure 1 in Galvanic Cells, the shorthand notation is

$$\text{Pt(s)}|\text{Cl}_2(\text{g}) \mid \text{Cl}^{-}(1\,\text{M}) \parallel \text{Fe}^{2+}(1\,\text{M}), \text{Fe}^{3+}(1\,\text{M}) \mid \text{Pt(s)}$$
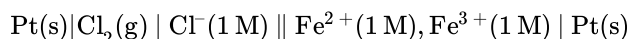
According to the conventions we have just developed, this corresponds to the cell reaction

$$2\,\text{Cl}^{-}(\text{aq}) + 2\,\text{Fe}^{3+}(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + 2\,\text{Fe}^{2+}(\text{aq})$$
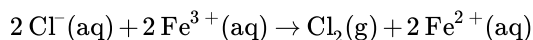
But this is the of the spontaneous cell reaction we described before (Eq. (2) from Galvanic Cells). Since the cell reaction is nonspontaneous, electrons will not be forced into the external circuit at the left-hand electrode, and they will not be withdrawn at the right. In fact the reverse will actually occur. Thus if a voltmeter is connected to this cell, its right-hand terminal will become more negative and its left-hand terminal will become more positive. This is shown in Figure 2 from Galvanic Cells.

In general, if a galvanic cell is connected to a voltmeter, the electrode connected to the negative terminal of the meter must be the anode. If our shorthand cell notation shows that electrode on the left, then the corresponding cell reaction must be spontaneous. Electrons will be released by the oxidation half-equation on the left and accepted by the reduction on the right. If, on the other hand, the voltmeter shows that the right-hand electrode is releasing electrons, then we must have written our shorthand notation *backwards*. This means that the reverse of the cell reaction obtained by our rules must actually be occurring, and it is that reverse reaction which is spontaneous. Thus by simply observing which electrode in the cell releases electrons and which accepts them, that is, by finding which electrode is negative and which positive, we can determine whether the cell reaction is spontaneous.