# Section 10.1: Galvanic Cells

Section 10.1 Questions, page 641

1. It is necessary to separate the oxidizing agent from the reducing agent in a galvanic cell because if the oxidizing agent and reducing agent in a galvanic cell were not separated, the chemical energy of the transferred electrons would be released into the environment as thermal energy and could not be harnessed to do useful work. When the oxidizing agent and reducing agent are separated, electron transfer can occur through a wire and produce electric current, which can be put to use operating various devices.

2. The two "connections" that are required before a galvanic cell can function are a wire to connect the two electrodes and a salt bridge to connect the two solutions.

3. The solution in a salt bridge should contain a soluble electrolyte, and it should not react with the electrodes or the electrolyte solutions in the cell.

4. (a), (c)
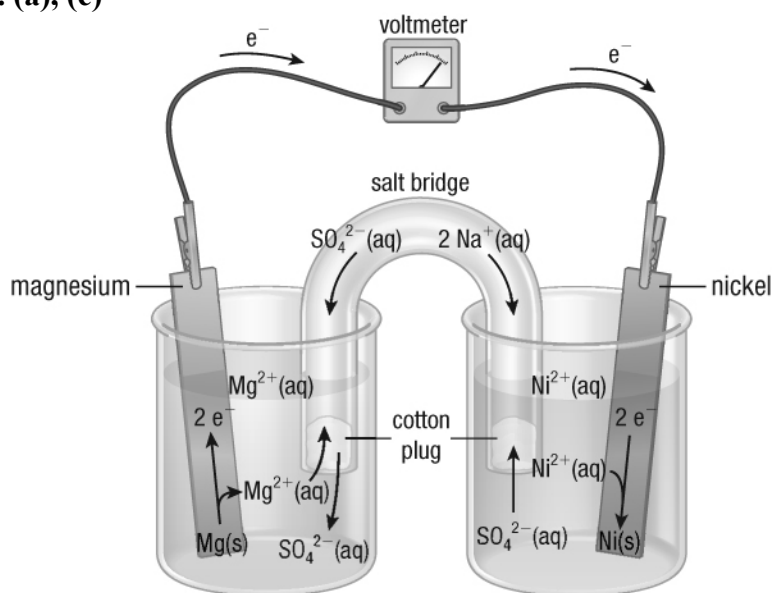
(b) anode: magnesium; cathode: nickel

(d) reducing agent: magnesium metal; oxidizing agent: nickel(II) ions

(e) The following half-reactions will occur:

Anode (oxidation) half-reaction equation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

Cathode (reduction) half-reaction equation: $Ni^{2+}(aq) + 2\ e^- \rightarrow Ni(s)$

(f) Net ionic equation: $Ni^{2+}(aq) + Mg(s) \rightarrow Ni(s) + Mg^{2+}(aq)$

(g) The mass of the magnesium electrode will become smaller and the mass of the nickel electrode will become larger as the cell operates.

5. (a) For the galvanic cell represented by $Mg(s)\ |\ Mg^{2+}(aq)\ ||\ Cu^+(aq)\ |\ Cu(s)$:

(i) stronger oxidizing agent: copper(I) ions; stronger reducing agent: magnesium metal

(ii) The following half reactions will occur:

Anode (oxidation) half-reaction equation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

Cathode (reduction) half-reaction equation: $Cu^+(aq) + e^- \rightarrow Cu(s)$

(iii) Net ionic equation: $Mg(s) + 2\ Cu^+(aq) \rightarrow Mg^{2+}(aq) + 2\ Cu(s)$

(b) For the galvanic cell represented by $Fe(s) | Fe^{2+}(aq) || Ag^+(aq) | Ag(s)$:

(i) stronger oxidizing agent: silver ions; stronger reducing agent: iron metal

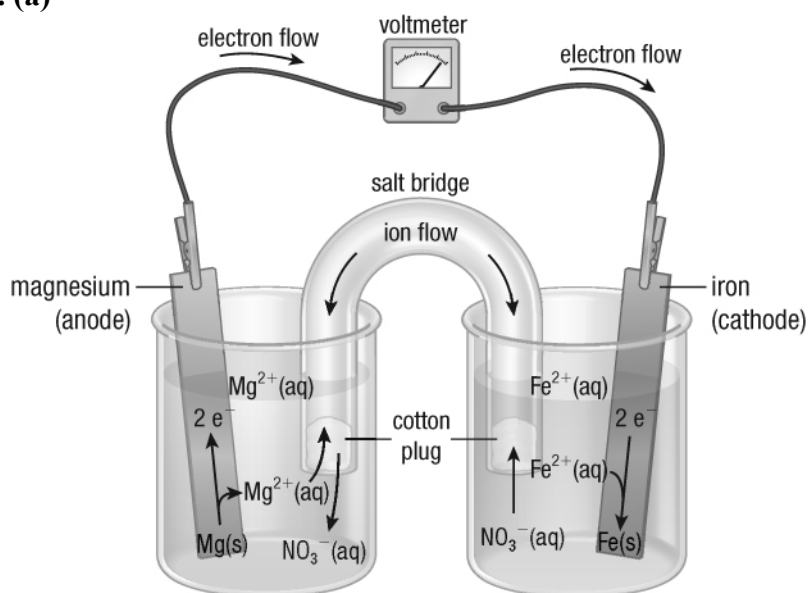**(ii)** The following half-reactions will occur:

Anode (oxidation) half-reaction equation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2\ e^-$

Cathode (reduction) half-reaction equation: $Ag^+(aq) + e^- \rightarrow Ag(s)$

(iii) Net ionic equation: $Fe(s) + 2\ Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2\ Ag(s)$

6. (a)

(b) The following half-reactions will occur:

Anode (oxidation) half-reaction equation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

Cathode (reduction) half-reaction equation: $Fe^{2+}(aq) + 2\ e^- \rightarrow Fe(s)$

(c) Net ionic equation: $Fe^{2+}(aq) + Mg(s) \rightarrow Fe(s) + Mg^{2+}(aq)$

(d) When the cell is functioning, I would expect to see the voltmeter indicating that the magnesium electrode has a negative charge and the iron electrode has a positive charge. Over time, I would also expect the magnesium electrode to become smaller and the iron electrode to become larger.

7. (a) The cell in Figure 7 has no half-cells and no salt bridge, whereas the zinc–copper cell in Figure 5 has both these features. The cell in Figure 7 has only one electrolyte solution: tap water.

(b) Anode (oxidation) half-reaction equation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

Cathode (reduction) half-reaction equation: $2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$

Net ionic equation: $Mg(s) + 2\ H_2O(l) \rightarrow Mg^{2+}(aq) + H_2(g) + 2\ OH^-(aq)$

(c) In this cell, copper acts as an inert electrode that provides a place to connect the wire to complete the circuit and as a surface on which the half-reactions can occur.

(d) The cell in Figure 7 works with tap water but not with distilled water because distilled water is pure water. Tap water also contains electrolytes, such as calcium and magnesium, which make it a better conductor of electricity than distilled water.