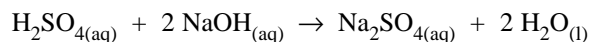
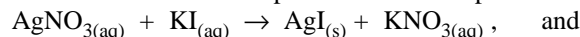


## UNIT 5 REVIEW

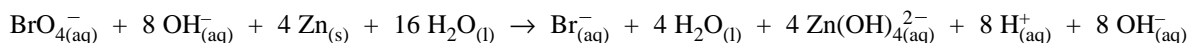
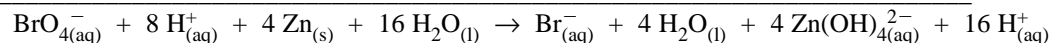
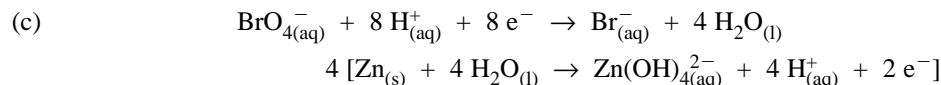
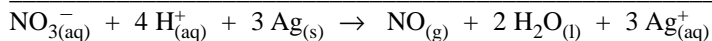
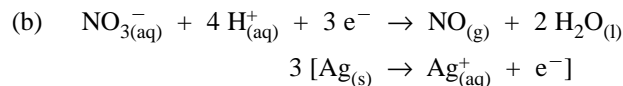
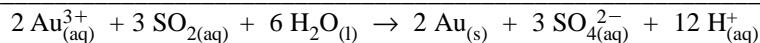
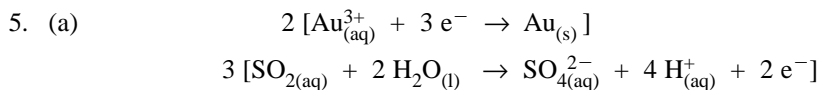
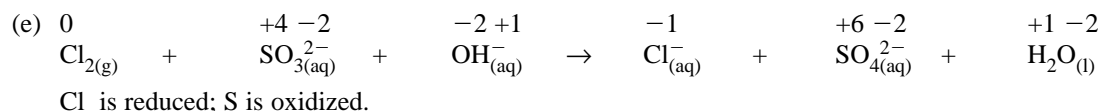
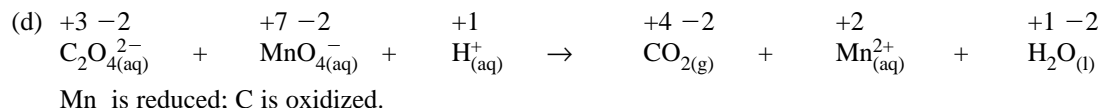
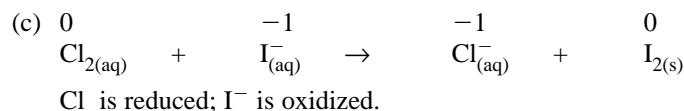
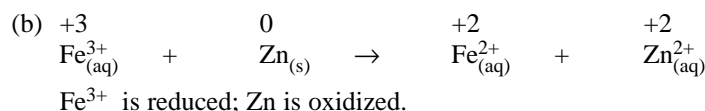
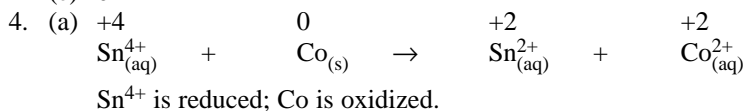
(Page 766)

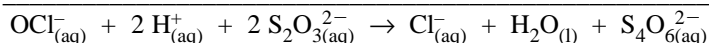
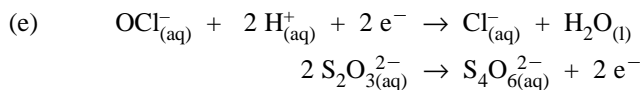
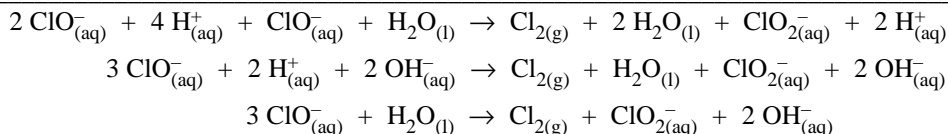
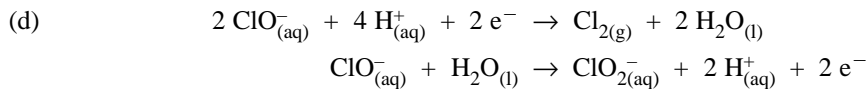
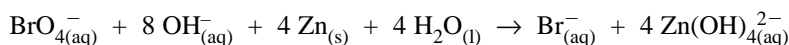
### Understanding Concepts

- (a) Reduction is a process in which electrons are gained; and the oxidation number of an atom decreases.  
(b) Oxidation is a process in which electrons are lost; and the oxidation number of an atom increases.  
(c) A redox reaction is one involving the transfer of electrons from one particle to another.
- Some chemical reactions are not redox reactions. They can be recognized by the fact that no entity changes in oxidation number. Most examples are double displacement reactions, such as:



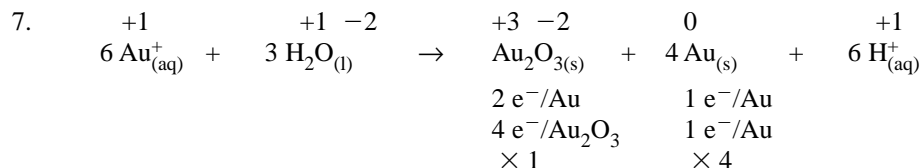
- (a) -2  
(b) +4  
(c) +6  
(d) +4  
(e) 0





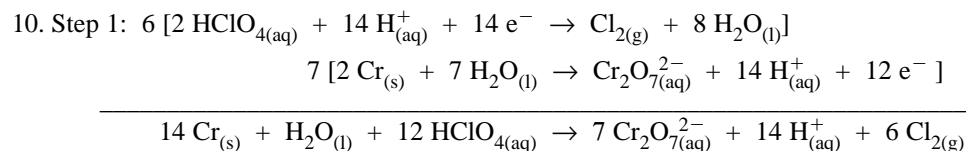
**Note:** Hypochlorite ion is commonly written both as  $\text{ClO}^-$ , and as  $\text{OCl}^-$ . It has been written in different ways in questions (d) and (e), to familiarize students with the use of both symbolisms.

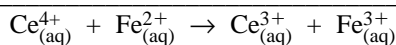
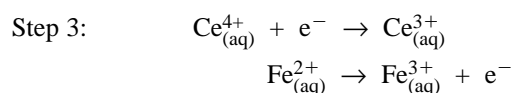
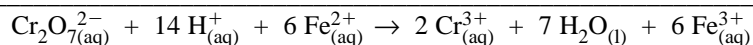
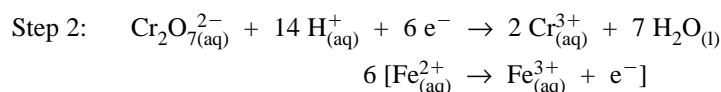
6. (a) Oxidation numbers:  $\text{Cu}^+$  is +1, Cu is 0, and  $\text{Cu}^{2+}$  is +2  
 (b)  $\text{Cu}_{(\text{aq})}^+ \rightarrow \text{Cu}_{(\text{aq})}^{2+} + \text{e}^-$   
 (c)  $\text{Cu}_{(\text{aq})}^+ + \text{e}^- \rightarrow \text{Cu}_{(\text{s})}$   
 (d)  $2 \text{Cu}_{(\text{aq})}^+ \rightarrow \text{Cu}_{(\text{aq})}^{2+} + \text{Cu}_{(\text{s})}$   
 (e) Three other ions that can behave as oxidizing or reducing agents are  $\text{Sn}_{(\text{aq})}^+$ ,  $\text{Cr}_{(\text{aq})}^{2+}$ , and  $\text{Fe}_{(\text{aq})}^{2+}$ .



**Note:** When particles of a single reactant are both oxidized and reduced (auto-oxidation, or self-oxidation–reduction), then “bookkeeping” for electron gain and loss balancing can be done on the product side of the equation, as shown above.

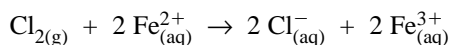
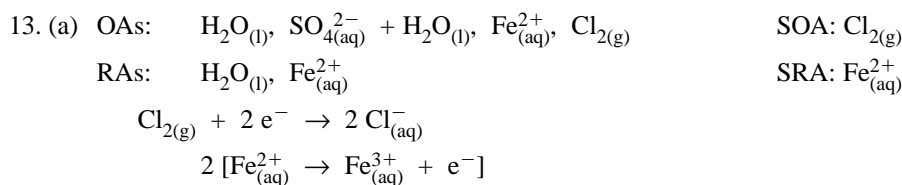
8. (a)  $\text{I}_{(\text{aq})}^- + 3 \text{H}_2\text{O}_{(\text{l})} \rightarrow \text{IO}_{3(\text{aq})}^- + 6 \text{H}_{(\text{aq})}^+ + 6 \text{e}^-$   
 (b)  $\text{Cl}_{2(\text{aq})} + 2 \text{e}^- \rightarrow 2 \text{Cl}_{(\text{aq})}^-$  [  $\times 3$ , to balance electrons ]  
 (c)  $\text{I}_{(\text{aq})}^- + 3 \text{H}_2\text{O}_{(\text{l})} + 3 \text{Cl}_{2(\text{aq})} \rightarrow \text{IO}_{3(\text{aq})}^- + 6 \text{H}_{(\text{aq})}^+ + 6 \text{Cl}_{(\text{aq})}^-$   
 9. (a)  $\text{Ni}_{(\text{s})} + 2 \text{H}_2\text{SO}_{4(\text{aq})} \rightarrow \text{NiSO}_{4(\text{aq})} + 2 \text{H}_2\text{O}_{(\text{l})} + \text{SO}_{2(\text{g})}$   
 (b)  $\text{I}_{2(\text{s})} + 10 \text{NO}_{3(\text{aq})}^- + 8 \text{H}_{(\text{aq})}^+ \rightarrow 2 \text{IO}_{3(\text{aq})}^- + 10 \text{NO}_{2(\text{s})} + 4 \text{H}_2\text{O}_{(\text{l})}$   
 (c)  $\text{Cr}_2\text{O}_{7(\text{aq})}^{2-} + 14 \text{H}_{(\text{aq})}^+ + 6 \text{Cl}_{(\text{aq})}^- \rightarrow 2 \text{Cr}_{(\text{aq})}^{3+} + 3 \text{Cl}_{2(\text{g})} + 7 \text{H}_2\text{O}_{(\text{l})}$   
 (d)  $4 \text{Zn}_{(\text{s})} + 5 \text{H}_2\text{SO}_{4(\text{aq})} \rightarrow 4 \text{ZnSO}_{4(\text{aq})} + \text{H}_2\text{S}_{(\text{g})} + 4 \text{H}_2\text{O}_{(\text{l})}$   
 (e)  $3 \text{I}_{2(\text{s})} + 6 \text{OH}_{(\text{aq})}^- \rightarrow 5 \text{I}_{(\text{aq})}^- + \text{IO}_{3(\text{aq})}^- + 3 \text{H}_2\text{O}_{(\text{l})}$



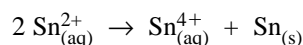
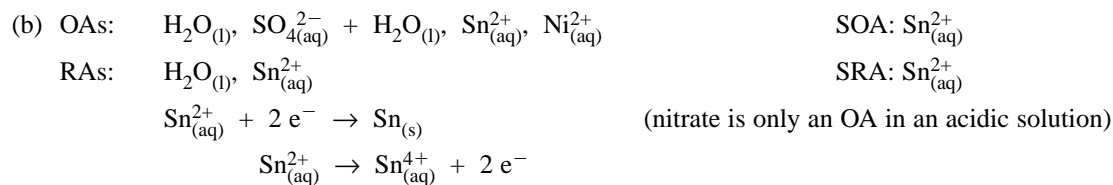


11. When a spontaneous redox reaction occurs, one might observe formation of a gas and/or solid, and/or a colour change, and/or an energy change.

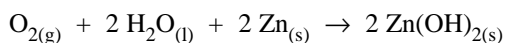
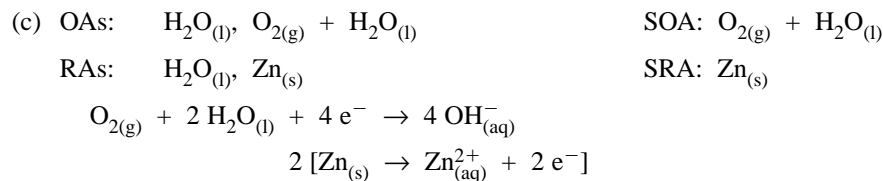
12. (a) nonspontaneous  
 (b) spontaneous  
 (c) nonspontaneous



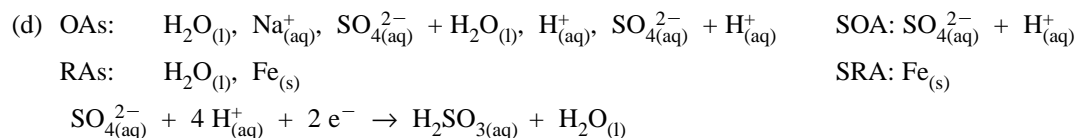
The strongest oxidizing agent present in the mixture is higher in the table of reduction potentials than is the strongest reducing agent, so the most probable oxidation–reduction reaction should be spontaneous.

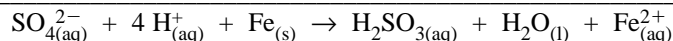
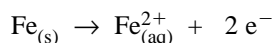


The strongest oxidizing agent present in the mixture is lower in the table of reduction potentials than is the strongest reducing agent, so the most probable oxidation–reduction reaction should be nonspontaneous.

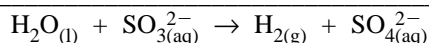
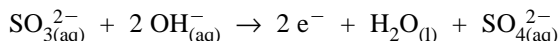
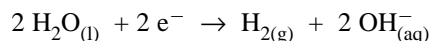
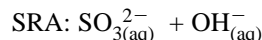
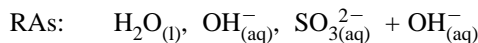
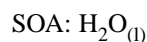
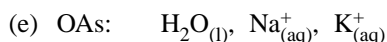


The strongest oxidizing agent present in the mixture is higher in the table of reduction potentials than is the strongest reducing agent, so the most probable oxidation–reduction reaction should be spontaneous.





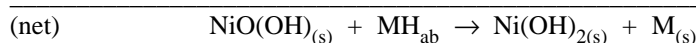
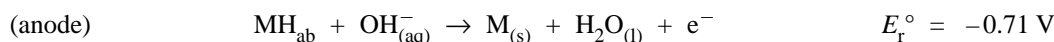
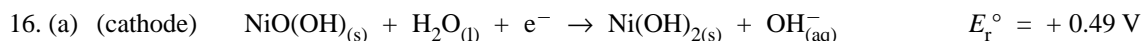
The strongest oxidizing agent present is higher in the table of reduction potentials than is the strongest reducing agent, so the most probable oxidation–reduction reaction should be spontaneous.



The strongest oxidizing agent present is higher in the table of reduction potentials than is the strongest reducing agent, so the most probable oxidation–reduction reaction should be spontaneous.

14. Two technological solutions to the problem of batteries going “dead” are secondary (rechargeable) cells, and fuel cells.

15. Technology preceded scientific explanations in the areas of reduction of metals from ores, and in the oxidation of metals, and in the development of electric cells to supply electric current.



(b)  $\Delta E^{\circ} = E_r^{\circ}(\text{cathode}) - E_r^{\circ}(\text{anode})$

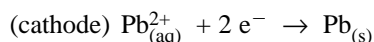
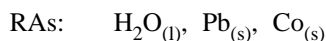
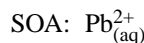
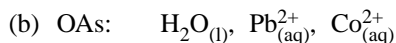
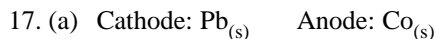
$= +0.49\text{ V} - (-0.71\text{ V})$

$\Delta E^{\circ} = +1.20\text{ V}$

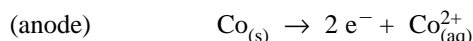
The cell potential of a NiMH battery is +1.20 V.

(c) Considerations for evaluating the NiMH battery include:

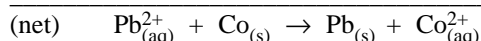
- Economic: NiMH cells are considerably more expensive than alkaline cells or NiCd rechargeables.
- Technological: NiMH cells contain more energy than other rechargeables, and can be recharged many more times before failing, and can deliver higher current (more power) than other types.
- Ecological/environmental: NiMH cells contain no heavy metals and fewer toxic materials than other rechargeables, making them less dangerous and also easier to recycle.



$E_r^{\circ} = -0.13\text{ V}$



$E_r^{\circ} = -0.28\text{ V}$

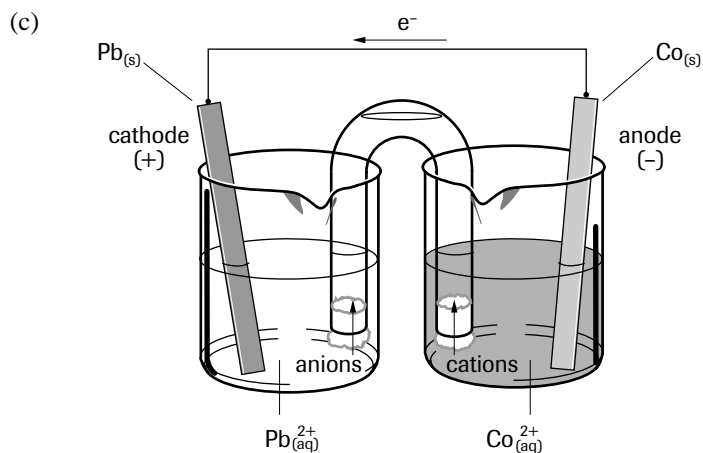


$\Delta E^{\circ} = E_r^{\circ}(\text{cathode}) - E_r^{\circ}(\text{anode})$

$= -0.13\text{ V} - (-0.28\text{ V})$

$\Delta E^{\circ} = +0.15\text{ V}$

The potential of a lead–cobalt standard cell is 0.15 V.



18.  $\Delta E^\circ = E_{\text{r}}^\circ (\text{cathode}) - E_{\text{r}}^\circ (\text{anode})$

(a)  $\Delta E^\circ = [-0.40 - (-0.91)] \text{ V}$

$\Delta E^\circ = 0.51 \text{ V}$

(b)  $\Delta E^\circ = [-0.13 - (-0.76)] \text{ V}$

$\Delta E^\circ = 0.63 \text{ V}$

(c)  $\Delta E^\circ = [+1.23 - (-0.28)] \text{ V}$

$\Delta E^\circ = 1.51 \text{ V}$

19. (a)  $\Delta E^\circ = E_{\text{r}}^\circ (\text{cathode}) - E_{\text{r}}^\circ (\text{anode})$

$E_{\text{r}}^\circ (\text{anode}) = E_{\text{r}}^\circ (\text{cathode}) - \Delta E^\circ$

$E_{\text{r}}^\circ (\text{anode}) = +0.34 \text{ V} - (+0.48 \text{ V})$  (Note that  $\text{Cu}^{2+}$  is reduced; therefore, Cu is the cathode.)

$E_{\text{r}}^\circ (\text{anode}) = -0.14 \text{ V}$

This answer matches the standard reduction potential of the  $\text{Sn}_{(\text{aq})}^{2+} + 2 \text{e}^- \rightarrow \text{Sn}_{(\text{s})}$  half-reaction. Therefore, X may be Sn.

(b) The negative value obtained for this reduction potential means that tin(II) ions are weaker oxidizing agents than hydrogen ions, under standard conditions.

20. The cadmium/cadmium(II) reduction potential is  $-0.40 \text{ V}$  and the cell potential is  $+1.85 \text{ V}$ . This means that the neodymium/neodymium(III) reduction potential must, therefore, either be  $1.85 \text{ V}$  higher, or lower, than  $-0.40 \text{ V}$ . If higher, neodymium is the cathode, and

$\Delta E^\circ = E_{\text{r}}^\circ (\text{cathode}) - E_{\text{r}}^\circ (\text{anode})$

$E_{\text{r}}^\circ (\text{cathode}) = \Delta E^\circ + E_{\text{r}}^\circ (\text{anode})$

$E_{\text{r}}^\circ (\text{cathode}) = +1.85 \text{ V} + (-0.40 \text{ V})$

$E_{\text{r}}^\circ (\text{cathode}) = +1.45 \text{ V}$

or if lower, neodymium is the anode, and

$E_{\text{r}}^\circ (\text{anode}) = E_{\text{r}}^\circ (\text{cathode}) - \Delta E^\circ$

$E_{\text{r}}^\circ (\text{anode}) = -0.40 \text{ V} - (+1.85 \text{ V})$

$E_{\text{r}}^\circ (\text{anode}) = -2.25 \text{ V}$

The standard reduction potential for the  $\text{Nd}_{(\text{aq})}^{3+} | \text{Nd}_{(\text{s})}$  half-cell is either  $+1.45 \text{ V}$ , or  $-2.25 \text{ V}$ .

21. Gold, silver, and platinum are relatively unreactive as reducing agents, and thus resist corrosion.

22. Sterling silver contains copper, which is a much stronger reducing agent than silver. In effect, the presence of the copper makes the alloy more reactive because the copper acts as a reducing agent in the presence of oxidizing agents such as oxygen.

23. Chromium plating gives steel a very shiny surface. Zinc plating produces a mottled surface, and tin plating a fairly dull surface. All three metals are easily oxidized in air but the oxide layer on each metal is generally very thin and adheres tightly to the metal. Chromium and zinc are both stronger reducing agents than iron, meaning that if the plating is scratched, they will react in preference to the iron. If a tin plating surface is broken, a spontaneous reaction occurs between the iron and the tin ions on the tin's surface. This means that the presence of the tin will then actually accelerate any corrosion of the steel underneath.
24. An impressed current can be used to prevent the corrosion of a buried steel pipe by applying a voltage that forces the steel to act as a cathode, thus preventing oxidation of iron atoms.
25. Corrosion often occurs where two different metals join because any moisture connecting the two metals sets up a tiny galvanic cell (like Volta's cell). Two different metals have different electric potentials, which creates an electric current in the presence of an electrolyte.
26. The main differences between galvanic and electrolytic cells in terms of their purpose is that galvanic cells are designed to produce electrical energy, while electrolytic cells are designed to produce chemical substances using electrical energy. The types of half-reactions that occur are essentially the same, except that they are spontaneous only in galvanic cells, and must be forced by an applied potential in electrolytic cells.
27. (a) (cathode)  $\text{K}_{(\text{l})}^{+} + \text{e}^{-} \rightarrow \text{K}_{(\text{l})}$  (potassium is liquid above 63°C)  
 (reaction with water)  $2 \text{K}_{(\text{s})} + 2 \text{H}_2\text{O}_{(\text{l})} \rightarrow \text{H}_{2(\text{g})} + 2 \text{K}_{(\text{aq})}^{+} + 2 \text{OH}_{(\text{aq})}^{-}$   
 (gas combustion)  $2 \text{H}_{2(\text{g})} + \text{O}_{2(\text{g})} \rightarrow 2 \text{H}_2\text{O}_{(\text{g})}$  (water forms as vapour in an open flame)
- (b) The gas burned with a violet flame because of the presence of traces of potassium ions.
28. (a) It was necessary to melt the compounds because only in molten state are the ions free to move to complete the charge transfer within the liquid.  
 (b) Electrolyzing aqueous solutions of the compounds will not work because water is a stronger oxidizing agent and would react preferentially to the metal ions.
29. (a) (cathode)  $2 [\text{Ca}_{(\text{l})}^{2+} + 2 \text{e}^{-} \rightarrow \text{Ca}_{(\text{s})}]$   
 (anode)  $2 \text{O}_{(\text{l})}^{2-} \rightarrow 4 \text{e}^{-} + \text{O}_{2(\text{g})}$   


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 (net)  $2 \text{Ca}_{(\text{l})}^{2+} + 2 \text{O}_{(\text{l})}^{2-} \rightarrow 2 \text{Ca}_{(\text{s})} + \text{O}_{2(\text{g})}$
- (b) (cathode)  $4 [\text{Na}_{(\text{l})}^{+} + \text{e}^{-} \rightarrow \text{Na}_{(\text{s})}]$   
 (anode)  $4 \text{OH}_{(\text{l})}^{-} \rightarrow 4 \text{e}^{-} + \text{O}_{2(\text{g})} + 2 \text{H}_2\text{O}_{(\text{g})}$   


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 (net)  $4 \text{Na}_{(\text{l})}^{+} + 4 \text{OH}_{(\text{l})}^{-} \rightarrow 4 \text{Na}_{(\text{s})} + \text{O}_{2(\text{g})} + 2 \text{H}_2\text{O}_{(\text{g})}$
- (c) (cathode)  $2 [\text{Mg}_{(\text{l})}^{2+} + 2 \text{e}^{-} \rightarrow \text{Mg}_{(\text{s})}]$   
 (anode)  $2 \text{O}_{(\text{l})}^{2-} \rightarrow 4 \text{e}^{-} + \text{O}_{2(\text{g})}$   

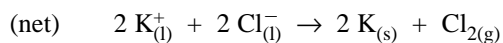
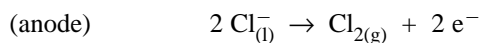
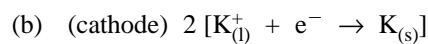
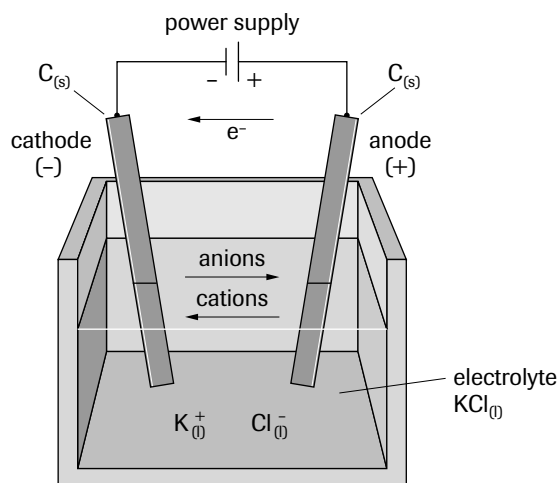

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 (net)  $2 \text{Mg}_{(\text{l})}^{2+} + 2 \text{O}_{(\text{l})}^{2-} \rightarrow 2 \text{Mg}_{(\text{s})} + \text{O}_{2(\text{g})}$
- (d) (cathode)  $2 [\text{Ba}_{(\text{l})}^{2+} + 2 \text{e}^{-} \rightarrow \text{Ba}_{(\text{s})}]$   
 (anode)  $4 \text{OH}_{(\text{l})}^{-} \rightarrow 4 \text{e}^{-} + \text{O}_{2(\text{g})} + 2 \text{H}_2\text{O}_{(\text{g})}$   

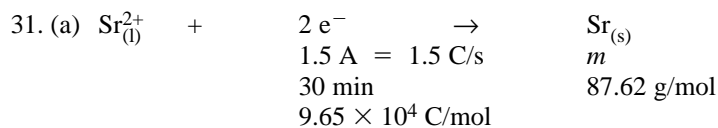

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 (net)  $2 \text{Ba}_{(\text{l})}^{2+} + 4 \text{OH}_{(\text{l})}^{-} \rightarrow 2 \text{Ba}_{(\text{s})} + \text{O}_{2(\text{g})} + 2 \text{H}_2\text{O}_{(\text{g})}$

30. (a)



(c) Our redox table has reduction potentials only for standard conditions: 1 mol/L electrolytes and SATP. The conditions of molten salt electrolysis are very different. Therefore, the reduction potentials would be very different.



$$t = 30 \cancel{\text{min}} \times \frac{60 \text{ s}}{1 \cancel{\text{min}}} = 1.8 \times 10^3 \text{ s}$$

$$n_{\text{e}^{-}} = \frac{q}{F}$$

$$n_{\text{e}^{-}} = \frac{It}{F}$$

$$= \frac{1.5 \frac{\text{C}}{\text{s}} \times 1.8 \times 10^3 \cancel{\text{s}}}{9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}$$

$$n_{\text{e}^{-}} = 0.0280 \text{ mol}$$

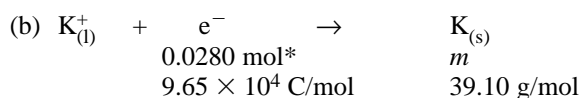
$$n_{\text{Sr}} = 0.0280 \cancel{\text{mol e}^{-}} \times \frac{1 \text{ mol Sr}}{2 \cancel{\text{mol e}^{-}}}$$

$$n_{\text{Sr}} = 0.0140 \text{ mol}$$

$$m_{\text{Sr}} = 0.0140 \cancel{\text{mol}} \times 87.62 \frac{\text{g}}{\cancel{\text{mol}}}$$

$$m_{\text{Sr}} = 1.23 \text{ g}$$

The mass of strontium produced is 1.23 g.



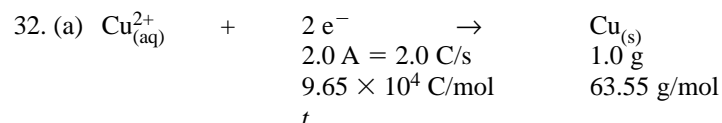
$$n_{\text{K}} = 0.0280 \text{ mol} \cancel{\text{e}^{-}} \times \frac{1 \text{ mol K}}{1 \cancel{\text{mol e}^{-}}} \quad (* \text{ the amount of charge is identical to that in part (a)})$$

$$n_{\text{K}} = 0.0280 \text{ mol}$$

$$m_{\text{K}} = 0.0280 \text{ mol} \times 39.10 \frac{\text{g}}{\text{mol}}$$

$$m_{\text{K}} = 1.09 \text{ g}$$

The mass of potassium produced is 1.09 g.



$$n_{\text{Cu}} = 1.0 \text{ g} \times \frac{1 \text{ mol}}{63.55 \text{ g}}$$

$$n_{\text{Cu}} = 0.016 \text{ mol}$$

$$n_{\text{e}^{-}} = 0.016 \text{ mol Cu} \times \frac{2 \text{ mol e}^{-}}{1 \cancel{\text{mol Cu}}}$$

$$n_{\text{e}^{-}} = 0.031 \text{ mol}$$

$$n_{\text{e}^{-}} = \frac{q}{F}$$

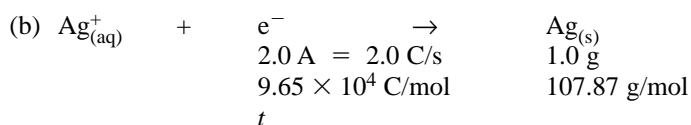
$$n_{\text{e}^{-}} = \frac{It}{F}$$

$$t = \frac{n_{\text{e}^{-}} F}{I}$$

$$= \frac{0.031 \cancel{\text{mol}} \times 9.65 \times 10^4 \frac{\text{C}}{\cancel{\text{mol}}}}{2.0 \frac{\text{C}}{\text{s}}}$$

$$t = 1.5 \times 10^3 \text{ s}$$

The time to plate 1.0 g of copper is  $1.5 \times 10^3 \text{ s}$  (or 0.42 h).



$$n_{\text{Ag}} = 1.0 \text{ g} \times \frac{1 \text{ mol}}{107.87 \text{ g}}$$

$$n_{\text{Ag}} = 0.0093 \text{ mol}$$

$$n_{\text{e}^{-}} = 0.0093 \text{ mol Ag} \times \frac{1 \text{ mol e}^{-}}{1 \cancel{\text{mol Ag}}}$$

$$n_{\text{e}^{-}} = 0.0093 \text{ mol}$$

$$n_{\text{e}^{-}} = \frac{q}{F}$$

$$n_{\text{e}^{-}} = \frac{It}{F}$$

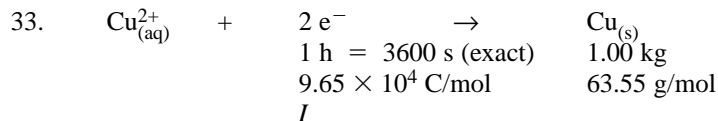
$$t = \frac{n_{\text{e}^{-}} F}{I}$$



$$= \frac{0.0093 \cancel{\text{mol}} \times 9.65 \times 10^4 \frac{\text{C}}{\cancel{\text{mol}}}}{2.0 \frac{\text{C}}{\text{s}}}$$

$$t = 4.5 \times 10^2 \text{ s}$$

The time to plate 1.0 g of silver is  $4.5 \times 10^2 \text{ s}$  (or 0.12 h).



$$n_{\text{Cu}} = 1.00 \text{ kg} \times \frac{1 \text{ mol}}{63.55 \text{ g}}$$

$$n_{\text{Cu}} = 0.0157 \text{ kmol} = 15.7 \text{ mol}$$

$$n_{\text{e}^-} = 15.7 \cancel{\text{mol Cu}} \times \frac{2 \text{ mol e}^-}{1 \cancel{\text{mol Cu}}}$$

$$n_{\text{e}^-} = 31.5 \text{ mol}$$

$$n_{\text{e}^-} = \frac{q}{F}$$

$$n_{\text{e}^-} = \frac{It}{F}$$

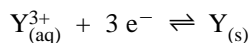
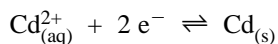
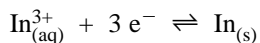
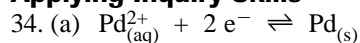
$$I = \frac{n_{\text{e}^-} F}{t}$$

$$= \frac{31.5 \cancel{\text{mol}} \times 9.65 \times 10^4 \frac{\text{C}}{\cancel{\text{mol}}}}{3600 \text{ s}}$$

$$I = 844 \text{ A}$$

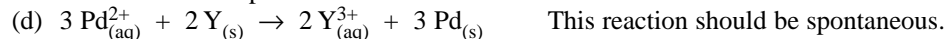
The average current needed to plate 1.00 kg of copper per hour is 844 A.

### Applying Inquiry Skills

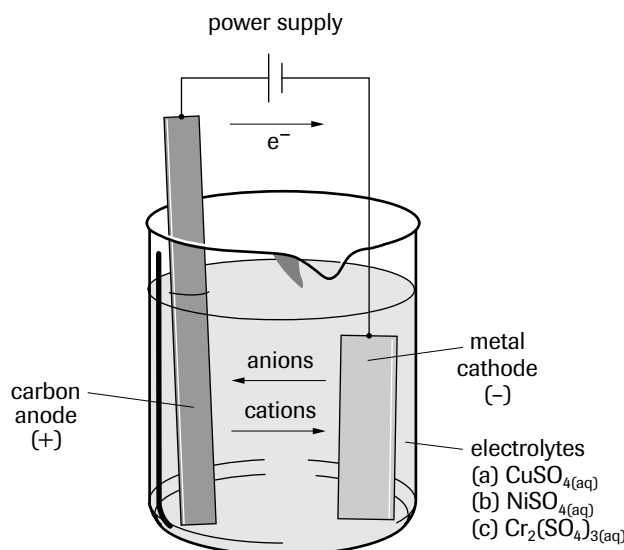


(b) A palladium(II) ion is the strongest oxidizing agent in the experiment because it attracts electrons better than any other metal ion.

(c) Yttrium metal is the strongest reducing agent in the experiment because it loses electrons more easily than any other metal atom present.



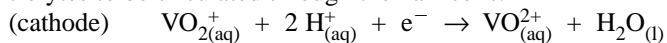
35. (a) A galvanized metal is made from the cathode of an electrolytic cell with a copper(II) sulfate electrolyte, and an inert carbon anode. After plating with copper, the electrolyte is replaced with nickel(II) sulfate. Finally, after plating with nickel, the electrolyte is replaced with chromium(III) sulfate.



- (b) The experimental variables include the current, the time of operation, and the volume and concentration of the electrolyte. Calculation begins with multiplying the surface area by thickness of plating to find the volume needed. The volume is then multiplied by density to find the mass of metal required. The mass is then divided by molar mass to find the amount (in moles) of metal required. The amount of metal is multiplied by the electron/atom mole ratio from the half-reaction equation to find the amount of electrons required. The amount of electrons is divided by the Faraday constant to find the charge that must be transferred. The transferred charge is related directly to both the current and the time.

### Making Connections

36. A vanadium redox battery (cell) consists of graphite electrodes, aqueous vanadium ions in dilute sulfuric acid, a plastic membrane that allows hydrogen ions to transfer between half-cells, and external tanks that store half-cell electrolytes to be circulated through the half-cells.



The vanadium redox battery uses the unique chemistry of vanadium to store electrical energy as chemical energy. It is like a fuel cell but rechargeable. The cell can be instantly recharged by replacing the electrolyte or recharged using an external power source (e.g., solar cells) to reverse the reactions. Because the electrolyte circulates between a half-cell and its storage tank, the electrolyte also acts as a coolant. Some of the advantages include the indefinite life of the solutions; high rates of recharge; and low cost, availability, and environmental friendliness of vanadium. The vanadium redox battery has been proposed for use as an electric power source for remote areas and as an energy storage medium for photovoltaic and wind energy systems.

37. Copper is better than aluminum for interconnects because copper has a lower electrical resistance and has fewer problems with electromigration. There will likely be a cost advantage because the interconnects can be thinner than those of aluminum. Copper interconnects are created by electroplating copper from a copper(II) ion solution. This is the reduction of copper in an electrolytic cell. Copper interconnects can also be made by electroless plating. In this method, copper(II) ion complexes are reduced to copper metal using a reducing agent such as formaldehyde. This is a spontaneous redox reaction in solution.
38. Road salt is often sodium chloride but can also be calcium chloride, magnesium chloride, and other salts. Ontario uses more salt per year than any other province (about 40% of the total Canadian use of salt). Road salt is used to melt snow and ice to improve driving conditions. The main advantages are reduced accident rates, fewer and less severe injuries, and fewer travelling delays. However, the use of road salt causes severe corrosion problems of vehicles, bridges, etc. The salt dissolved in water is a very good electrolyte and greatly improves the charge transfer in electrochemical cells causing an accelerated rate of corrosion. Salt is also potentially harmful to aquatic organisms and may also contaminate ground water. There are several alternatives to the current practice of using large quantities of solid sodium chloride. One alternative is calcium magnesium acetate, which has less ability in promoting corrosion and is less harmful to aquatic organisms, but it is relatively expensive. One of the most promising alternatives is to

treat the roads with a brine solution before significant icing occurs. This uses a lot less salt but requires very good weather forecasting to predict likely trouble areas.

### Extensions

39. To “boost” a dead car battery safely, the cables must be connected in a specific order.

- (1) The first connection is to clamp a positive (red) cable end to the positive terminal of the good battery.
- (2) The second connection is to clamp the other end of the positive cable to the positive terminal of the dead battery.
- (3) The third connection is to clamp a negative (black) cable end to the negative of the good battery — by connecting the cable clamp to any metal bracket or object that is part of the car body/frame structure, in a location well away from the battery location.
- (4) The fourth and last connection is made by clamping the other negative end of the cable to the negative of the dead battery — again, by connecting the cable clamp to any metal bracket or object that is part of the car structure, well away from the battery location.

The final two connections to “ground” (negative terminal connections) are made at a distance from both batteries to reduce the risk of possible explosion of battery gases (hydrogen and oxygen). Such explosions can be caused by sparking as the last connection (and later, the first disconnection) is made, if one is foolish enough to make and break this connection directly to a negative battery terminal!

40. (a) Proton transfer reactions are usually thought of as equilibrium situations, where the extent of the reaction is quite variable and is the primary concern. Electron transfer reactions tend to occur either to a negligible extent, or to such an extent they are considered complete. Both types occur readily in aqueous solution. Most proton transfer reactions primarily involve molecular compounds or organic ions. Electron transfer reactions commonly involve either molecules or ions. Many oxidizing agents react subject to the activity of hydrogen ions, in situations that are really both types of reactions.
- (b) Neutralization reactions in solution release heat as water molecules from which is evidence for energy change during proton transfer from hydronium to hydroxide ions. Combustion reactions are electron transfer reactions, and commonly are very exothermic.
- (c) Nitric acid is the classic case of a material that is both a strong acid and a strong oxidizing agent. In a reaction with barium hydroxide solution, pH testing and a gas collection and a glowing splint test for oxygen could be done to establish which products predominate.