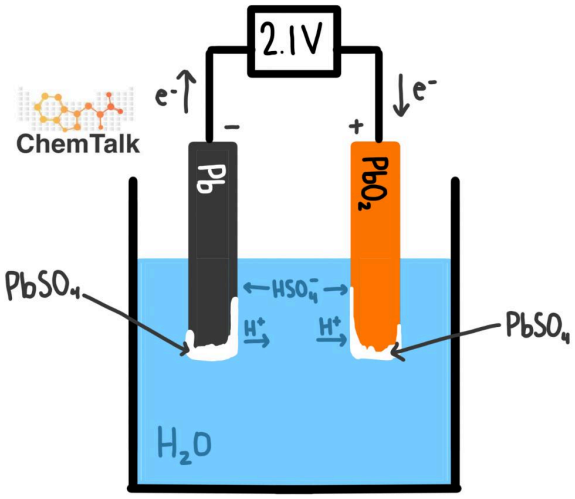
Batteries

This worksheet will cover batteries (vessels for redox reactions that result in an electric current). It will discuss lead-acid, nickel-cadmium, and fuel cell batteries. As you progress through the worksheet, you will develop the skills necessary to distinguish primary and secondary batteries, analyze reactions occurring in different batteries, and understand the importance of specific battery components.

Practice Problems:

1. What is the main difference between secondary batteries and primary batteries?

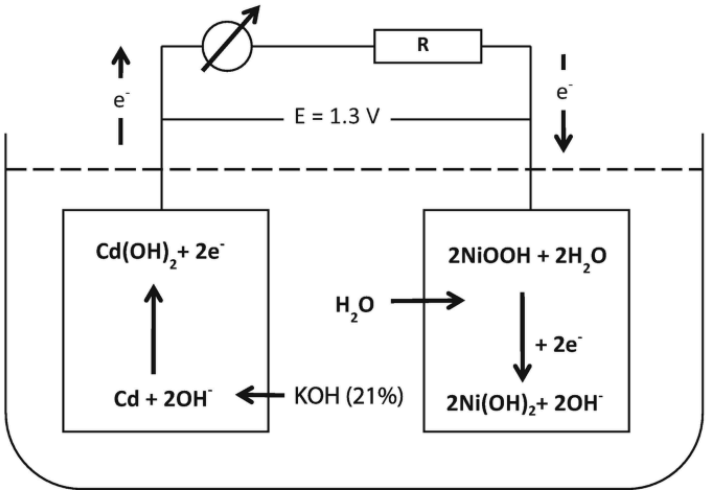
2. What is being oxidized in the forward reaction of lead-acid batteries?



3. What is being reduced in the forward reaction of lead-acid batteries?

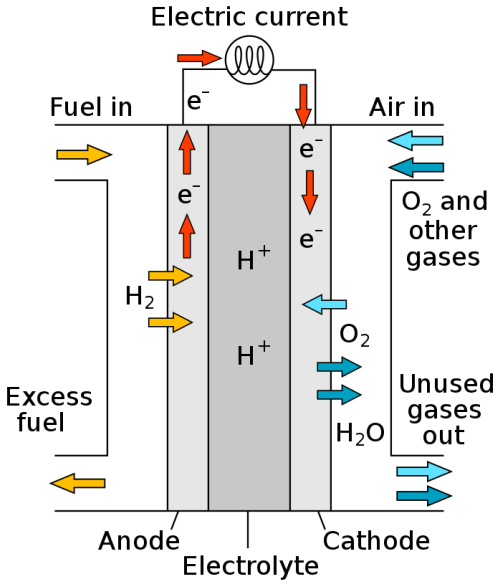
4. Why are lead-acid batteries also electrolytic cells (why do they require an input of electric current)?

5. In nickel-cadmium batteries, which products become a solid precipitate, and why do you need to apply a current to the battery?



Citation: https://link.springer.com/chapter/10.1007/978-3-030-57269-3_4

6. Describe how a fuel cell works. What are the redox reactions happening in a fuel cell?



Unit 16: Electrochemistry

ANSWER KEY

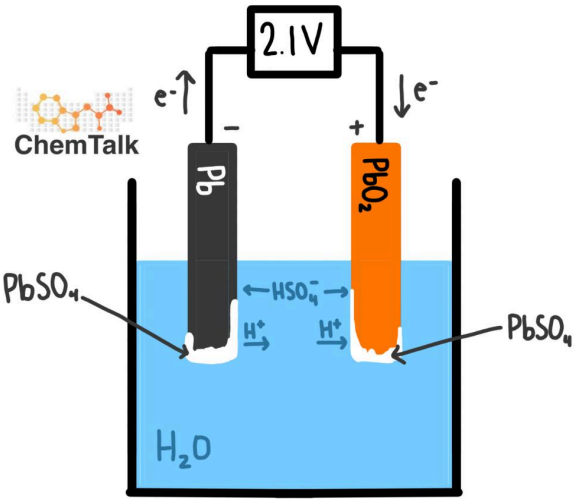
Batteries

Practice Problems:

1. What is the main difference between secondary batteries and primary batteries?

Primary batteries are single-use and disposable, like AA and AAA batteries in common household electronics. The redox reaction in primary batteries is irreversible, so the batteries cannot recharge. Secondary batteries, on the other hand, are rechargeable, since the redox reaction can move in the forward and reverse directions.

2. What is being oxidized in the forward reaction of lead-acid batteries?



Reaction of lead-acid batteries:

$$Pb\ (s) + PbO_2\ (s) + 2HSO_4^- + H^+ \rightleftharpoons 2PbSO_4\ (s) + 2H_2O$$

The neutral lead at the anode is being oxidized into lead (II) sulfate ($PbSO_4$). Since the oxidation state is increasing from 0 (neutral, free element) to +2 (Pb^{+2}), the lead atom is losing electrons and becoming more positive, indicating its oxidation.

3. What is being reduced in the forward reaction of lead-acid batteries?

Reaction of lead-acid batteries:

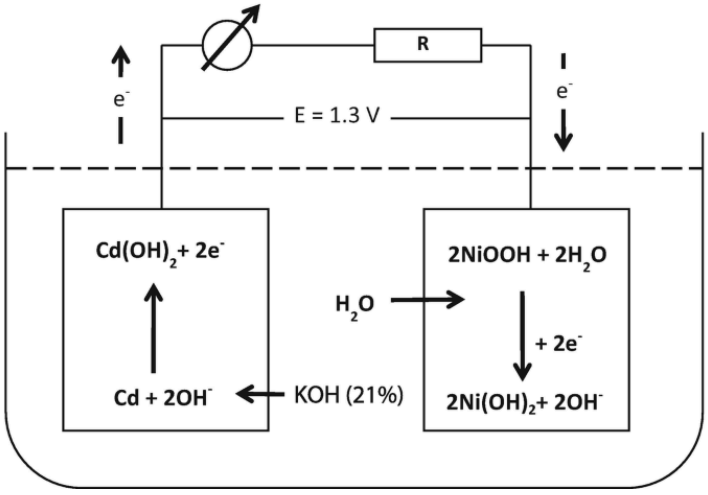
$$Pb\ (s) + PbO_2\ (s) + 2HSO_4^- + H^+ \rightleftharpoons 2PbSO_4\ (s) + 2H_2O$$

The lead (IV) oxide (PbO_2) at the cathode is reduced into $PbSO_4$. Since the oxidation is decreasing from +4 (Pb^{+4}) to +2 (Pb^{+2}), the lead atom in PbO_2 is gaining electrons and becoming more negative, indicating its reduction.

4. Why are lead-acid batteries also electrolytic cells (why do they require an input of electric current)?

Electrolytic cells are electrochemical cells that convert electrical energy from an external source into chemical energy (a non-spontaneous reaction). They require an input of energy. Lead-acid batteries produce $PbSO_4$, which is a solid precipitate. This product may be reconverted into elemental lead and PbO_2 with an input of electric current. Therefore, due to this reaction, lead acid batteries are also electrolytic cells.

5. In nickel-cadmium batteries, which products become a solid precipitate, and why do you need to apply a current to the battery?



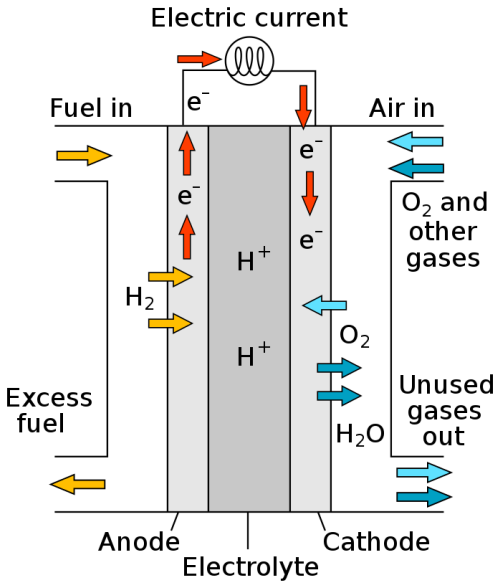
Citation: https://link.springer.com/chapter/10.1007/978-3-030-57269-3_4

Reaction of nickel-cadmium batteries:

$$Cd\ (s) + 2NiO(OH)\ (s) + 2H_2O \rightleftharpoons 2Ni(OH)_2\ (s) + Cd(OH)_2\ (s)$$

In nickel-cadmium batteries, the hydroxide products precipitate ($Ni(OH)_2$ & $Cd(OH)_2$) on the surface of the electrodes, but we can reverse the redox reaction with the application of an electric current. Thus, you need the input of energy to the battery to make it rechargeable.

6. Describe how a fuel cell works. What are the redox reactions happening in a fuel cell?



Fuel cells function like galvanic cells and require constant inputs of hydrogen and oxygen gas, with an output of H_2O (g). The electrolyte catalyzes a paired conversion of H_2 into H^+ and oxygen gas into O^{-2}, producing gaseous water. The reactions occurring in the cell are as written:

$$O_2\,(g) + 4e^- \rightarrow 2O^{-2}$$
$$2H_2\,(g) \rightarrow 4H^+ + 4e^-$$
$$O_2\,(g) + 2H_2\,(g) \rightarrow 4H^+ + 2O^{-2} \rightarrow 2H_2O\,(g)$$

Oxygen gas is being reduced, since it is gaining electrons. Hydrogen gas is being oxidized, since it is losing electrons.