B13 · Batteries, fuel cells and chlorine for water

Completed notes

Goals: Explain batteries as voltaic cells, and primary and rechargeable cells. Write the reactions and find the voltage of a lead-acid cell, an alkaline cell and a hydrogen fuel cell. Explain how chlorine is made for water, and judge a cell for a purpose using evidence.

Slide 1 of 10 · A battery is a voltaic cell

A battery is **one or more** voltaic cells. A voltaic cell has an anode, a cathode and an **electrolyte**, which is a liquid with dissolved ions.

When cells are joined in a row, their voltages **add**. So more cells in a row give **more** voltage.

A **primary** cell is used once, because its reaction cannot easily be reversed. A **secondary** cell can be recharged.

Recharging runs the cell reaction **backward** with a power supply. During recharge the battery acts as an **electrolytic** cell.

Example. A phone, a car and a watch each use stored chemical energy as electricity. What is a battery? How does a primary cell differ from a secondary cell?

1. A battery is one or more voltaic cells. Cells joined in a row add their voltages.

2. A primary cell is used once, because its reaction cannot easily be reversed.

3. A secondary cell can be recharged by running the reaction backward with a power supply.

Try it. A rechargeable battery is being recharged by a power supply. What kind of cell is it acting as? What happens to the electrode that was oxidized during discharge?

1. It acts as an electrolytic cell, because a power supply forces the reverse reaction.

2. The electrode that was oxidized is now reduced, so the anode and cathode names swap.

Slide 2 of 10 · The lead-acid battery

Two lead rows match this cell: E°(PbO_2/$PbSO_4$) = +1.69 V and E°($PbSO_4$/Pb) = **−0.36 V**. In a voltaic cell the **higher** value is the cathode.

The higher row runs as written at the **cathode**. The lower row is reversed to Pb + SO_4^{2-} → $PbSO_4$ + $2e^-$ at the **anode**, but its E° value stays **as printed**.

Both halves move **2** electrons, so the electrons cancel. The cell reaction is PbO_2 + Pb + $2SO_4^{2-}$ + $4H^+$ → **$2PbSO_4$ + $2H_2O$**. The acid gets **weaker**.

E°cell = E°cathode − E°anode = 1.69 − (−0.36) = **+2.05 V**. Six cells in a row give 6 × 2.05 = **12.30 V**, or 12.3 V to three digits, enough for a car.

Example. A car battery is rechargeable and gives large currents. Given E°(PbO_2/$PbSO_4$) = +1.69 V and E°($PbSO_4$/Pb) = −0.36 V, write the cell reaction, find E°cell for one cell and find the voltage of six cells in a row.

1. Cathode (higher value): PbO_2 + SO_4^{2-} + $4H^+$ + $2e^-$ → $PbSO_4$ + $2H_2O$

2. Anode (lower value, reversed): Pb + SO_4^{2-} → $PbSO_4$ + $2e^-$

3. Add the halves. The $2e^-$ cancel: PbO_2 + Pb + $2SO_4^{2-}$ + $4H^+$ → $2PbSO_4$ + $2H_2O$

4. Check each side: Pb 2, S 2, O 10, H 4, charge 0.

5. E°cell = 1.69 − (−0.36) = +2.05 V

6. Six cells in a row: 6 × 2.05 = 12.30 V, or 12.3 V to three digits

Try it. Given E°(PbO_2/$PbSO_4$) = +1.69 V and E°($PbSO_4$/Pb) = −0.36 V. Which row belongs to the anode? What is the voltage of a battery of four of these cells in a row?

1. The lower value, −0.36 V, belongs to the anode, so the $PbSO_4$/Pb row is reversed there.

2. One cell: E°cell = 1.69 − (−0.36) = +2.05 V.

3. Four cells in a row: 4 × 2.05 = 8.20 V.

Slide 3 of 10 · A primary cell: the alkaline cell

An alkaline cell is a **primary** cell. Its anode is **zinc** powder and its cathode is manganese(IV) oxide in a basic paste.

When half-reactions are **given**, use them and do not look in the Data Booklet. Zinc is oxidized: $Zn + 2OH^- \rightarrow ZnO + H_2O + \mathbf{2e^-}$.

Manganese is reduced: $2MnO_2 + H_2O + 2e^- \rightarrow \mathbf{Mn_2O_3} + 2OH^-$. Add the halves and cancel the electrons, water and hydroxide ions: $Zn + 2MnO_2 \rightarrow ZnO + Mn_2O_3$.

The zinc is used up, and the reaction cannot easily be reversed. Used cells hold metals that harm soil and water, so they should be **recycled**. A diploma answer gives a benefit and a **disposal** concern.

Example. A flashlight uses an alkaline cell that is not recharged. The given half-reactions are $Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$ and $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$. Find the overall reaction and say what happens to the zinc.

1. Both halves have $2e^-$, so the electrons cancel when we add.

2. Left: $Zn + 2OH^- + 2MnO_2 + H_2O$. Right: $ZnO + H_2O + Mn_2O_3 + 2OH^-$.

3. Cancel H_2O and OH^- from both sides: $Zn + 2MnO_2 \rightarrow ZnO + Mn_2O_3$

4. Check each side: Zn 1, Mn 2, O 4, charge 0.

5. Zinc is oxidized and is used up.

Try it. In the alkaline cell reaction $Zn + 2MnO_2 \rightarrow ZnO + Mn_2O_3$, find the oxidation number of zinc and of manganese on each side. What is oxidized and what is reduced?

1. Zinc goes from 0 in Zn to +2 in ZnO, so zinc is oxidized.

2. Manganese goes from +4 in MnO_2 to +3 in Mn_2O_3, so MnO_2 is reduced.

Slide 4 of 10 · Your turn: a lead-acid battery

Cells in a row **add** their voltages. The number of cells is the voltage needed divided by the **voltage of one cell**.

You cannot use part of a cell, so the number of cells must be a **whole number**. Follow the rounding rule in the question.

Oxidation happens at the anode. In $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$, the **lead** electrode is the anode. Solid $PbSO_4$ stays on it, so its mass goes **up**.

At the cathode, PbO_2 is **reduced** and solid $PbSO_4$ also forms there. So on discharge both electrodes **gain** mass, and the acid gets weaker.

Example. Given lead-acid cells in a row, each with $E°_{cell} = +2.05$ V. The discharge half-reactions are $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$ and $PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$. A 6.15 V battery is needed. How many cells are needed? Which electrode is the anode, and does its mass rise or fall?

1. Number of cells = 6.15 V ÷ 2.05 V per cell = 3.00, so 3 cells.

2. Lead loses electrons in $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$, so the lead electrode is the anode.

3. $PbSO_4$ is a solid that forms on the lead, so the anode mass rises.

Try it. A battery must give 10.25 V from lead-acid cells of 2.05 V each. How many cells are needed? Does the cathode mass rise or fall on discharge, and why?

1. Number of cells = 10.25 V ÷ 2.05 V per cell = 5.00, so 5 cells.

2. PbO_2 is reduced to solid $PbSO_4$, which forms on the cathode, so the cathode mass rises.

Slide 5 of 10 · Method: how to find how many cells make a battery

Step 1. Find the voltage of **one cell**. Use E°cell.

Step 2. **Divide** the voltage needed by the voltage of one cell.

Step 3. Round **up** to a whole number of cells, unless the question gives a rule.

Step 4. Multiply **back** to check the total voltage.

Example. Use the method on lead-acid cells of E°cell = +2.05 V. A 6.15 V battery is needed. How many cells?

1. Step 1: one cell gives +2.05 V.

2. Step 2: 6.15 ÷ 2.05 = 3.00.

3. Step 3: 3 cells.

4. Step 4: 3 × 2.05 = 6.15 V, which is at least 6.15 V.

Try it. Use the method on lead-acid cells of 2.05 V each. A battery must give at least 8.2 V. How many cells?

1. 8.2 ÷ 2.05 = 4.0, so 4 cells.

2. Check: 4 × 2.05 = 8.20 V, which reaches 8.2 V.

Slide 6 of 10 · The hydrogen fuel cell

A fuel cell is a voltaic cell that is **fed** its reactants all the time. A battery stores a fixed amount and runs down.

At the anode, hydrogen is oxidized: $2H_2 \rightarrow \mathbf{4H^+} + 4e^-$. At the cathode, oxygen is reduced: $O_2 + 4H^+ + 4e^- \rightarrow \mathbf{2H_2O}$.

Both halves move **4** electrons, so the electrons cancel. The overall reaction is $\mathbf{2H_2 + O_2 \rightarrow 2H_2O}$. $E°cell = 1.23 - 0.00 = \mathbf{+1.23\ V}$.

Do not multiply E° by a coefficient used to balance electrons. To judge a claim of no emissions, you need **evidence** about how the hydrogen is made and moved.

Example. A fuel cell makes electricity from hydrogen and oxygen that are fed to it all the time. Given $2H_2 \rightarrow 4H^+ + 4e^-$ and $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$. Also given $E°(O_2/H_2O) = +1.23$ V and $E°(H^+/H_2) = 0.00$ V. Write the overall reaction and find E°cell.

1. Both halves have $4e^-$, so the electrons cancel.

2. Add: $2H_2 + O_2 + 4H^+ \rightarrow 4H^+ + 2H_2O$. Cancel $4H^+$: $2H_2 + O_2 \rightarrow 2H_2O$

3. Check each side: H 4, O 2, charge 0.

4. $E°cell = 1.23 - 0.00 = +1.23$ V. The value is not multiplied by anything.

Try it. Each cell of a hydrogen fuel cell stack gives $E°cell = +1.23$ V. A stack has three cells in a row. What is the stack voltage? Why does a fuel cell keep running when a battery runs down?

1. Cells in a row add their voltages: $3 \times 1.23 = 3.69$ V.

2. A fuel cell is fed hydrogen and oxygen all the time, so its reactants are not used up.

Slide 7 of 10 · Chlorine for a swimming pool

In salt water, NaCl(aq), a current makes one gas at each electrode. At the cathode, **water** is reduced: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$.

At the anode, chloride is oxidized: $2Cl^- \rightarrow \mathbf{Cl_2} + 2e^-$. The +1.23 V water row is for acid, so the table gives only a first **prediction** that water is oxidized. Chlorine forms in practice. This is the chloride **anomaly**.

An oxidation number is the charge an atom would have if all shared electrons went to the atom that attracts them **more**. In $Cl_2 + 2OH^- \rightarrow Cl^- + OCl^- + H_2O$, chlorine goes from **0** to −1 and to **+1**.

The same element is both oxidized and reduced, which is called a **disproportionation**. Hypochlorite, OCl^-, cleans water because it is an **oxidizing** agent.

Example. Salt-water pools make chlorine by electrolysis of NaCl(aq). Name the gas formed at each electrode. Then use oxidation numbers to show that $Cl_2 + 2OH^- \rightarrow Cl^- + OCl^- + H_2O$ is a disproportionation.

1. Cathode: hydrogen gas forms. Anode: chlorine gas forms.

2. Chlorine in Cl_2 is 0. In Cl^- it is −1, so that chlorine is reduced.

3. In OCl^-, oxygen is −2 and the ion charge is 1−, so chlorine is +1 and is oxidized.

4. The same element is both oxidized and reduced, so this is a disproportionation.

Try it. Given $Br_2 + 2OH^- \rightarrow Br^- + OBr^- + H_2O$, with oxygen −2 in OBr^-. Find the oxidation number of bromine in each particle that contains it. Is this a disproportionation?

1. Bromine in Br_2 is 0. In Br^- it is −1. In OBr^- it is +1, because $-2 + (+1) = -1$.

2. Bromine is both reduced (0 to −1) and oxidized (0 to +1), so it is a disproportionation.

Slide 8 of 10 · Choose a cell for a purpose

To choose a cell, start with the **job**: what must the cell do? Then compare cells using evidence.

Voltage, mass, how long it runs, recharging, cost and disposal are each **separate** evidence. A higher voltage alone does not show a longer **life** or a lower cost.

A primary cell is cheap and keeps its charge for years. A rechargeable cell costs more at the start but can be used **many times**. A **fuel** cell runs as long as fuel keeps coming.

There is no best cell for every job. Justify a choice by naming one **advantage** and one **limit** of the type you choose.

Example. A company wants a cell for a flashlight that is used rarely and stored for years. Compare a primary cell, a rechargeable cell and a fuel cell. Which one would you choose? Give one advantage and one limit.

1. Choose a primary cell, such as an alkaline cell.

2. Advantage: it is cheap and keeps its charge in storage.

3. Limit: it is used once, and the used cell holds metals that must be recycled.

4. A rechargeable cell costs more at the start. A fuel cell needs fuel that has to be made and stored.

Try it. A company wants a cell for a laptop that is used every day. Which type would you choose? Give one advantage and one limit.

1. Choose a rechargeable cell.

2. Advantage: it can be used many times, so there is less waste.

3. Limit: it costs more at the start, and the battery still needs recycling at the end.

Slide 9 of 10 · Method: how to choose a cell for a purpose

Step 1. State what the **job** needs.

Step 2. List the **facts** about each cell: voltage, mass, life, recharging and cost.

Step 3. Match each need to a fact. Use **evidence**, not the highest voltage alone.

Step 4. Note disposal and safety. Used cells contain **metals**.

Step 5. Recommend one cell with a **reason** and one limit.

Example. Use the method on a flashlight that is used rarely and kept in a drawer for years. Which type of cell suits it?

1. Step 1: the job needs low cost and a charge that lasts in storage.

2. Steps 2 and 3: a primary cell is cheap and keeps its charge for years.

3. Step 4: the used cell holds zinc and manganese compounds, so recycle it.

4. Step 5: choose the primary cell. The limit is that it cannot be recharged.

Try it. Use the method on a phone that is used every day. Which type of cell suits it?

1. The job needs many uses at a low cost per use.

2. A rechargeable cell costs more at the start but can be recharged many times.

3. Recycle the old cell. Limit: it costs more at the start.

Slide 10 of 10 · Extracting metals long ago

Redox is much older than batteries. Early peoples heated copper ore with **charcoal** and got copper metal: $2CuO + C \rightarrow 2Cu + CO_2$.

Copper goes from **+2** to 0, so copper is reduced. Carbon goes from 0 to **+4**, so carbon is oxidized.

The reducing agent is the one that is **oxidized**. Its oxidation number **goes up**. Here the reducing agent is **carbon**.

Trial and error gave the method first, and the redox model explained **why** it works. Metals low on the table, like **aluminium**, need electrolysis instead.

Example. Early peoples heated copper(II) oxide with charcoal and got copper. Given $2CuO + C \rightarrow 2Cu + CO_2$, find the oxidation number of copper and of carbon on each side. Name what is oxidized and reduced, and the reducing agent. Check the atoms and charge.

1. Copper goes from +2 in CuO to 0 in Cu, so copper is reduced.

2. Carbon goes from 0 in C to +4 in CO_2, so carbon is oxidized.

3. Carbon gave electrons to copper(II), so carbon is the reducing agent.

4. Check each side: Cu 2, O 2, C 1, charge 0.

Try it. Iron(III) oxide is heated with carbon monoxide: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$. Find the oxidation number of iron and of carbon on each side. Name what is oxidized and the reducing agent.

1. Iron goes from +3 in Fe_2O_3 to 0 in Fe, so iron is reduced.

2. Carbon goes from +2 in CO to +4 in CO_2, so carbon is oxidized.

3. Carbon monoxide is oxidized, so it is the reducing agent.

4. Check: Fe 2, C 3, O 6 on each side. Electrons: 2Fe gain 6 and 3C lose 6.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 1-2 of 8

1. Explain the difference between a primary cell and a secondary cell, and say what happens during recharging.

A primary cell is used once. A secondary cell can be recharged. During recharging a power supply reverses the cell reaction, so the battery acts as an electrolytic cell.

2. Given these Data Booklet table rows for the lead-acid cell: $PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$ (+1.69 V) and $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$ (−0.36 V). Find E°cell for one cell and the voltage of a battery of six cells.

E°cell = 1.69 − (−0.36) = +2.05 V. Six cells in a row: 6 × 2.05 = 12.30 V, or 12.3 V to three digits.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 3-4 of 8

3. Given the alkaline cell half-reactions $Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$ and $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$. Name what is oxidized and what is reduced. Then write the overall reaction.

Zn is oxidized and MnO_2 is reduced. Overall: $Zn + 2MnO_2 \rightarrow ZnO + Mn_2O_3$.

4. Given a battery that must give 8.2 V from lead-acid cells of 2.05 V each. Find the number of cells and say which electrode gains mass on discharge.

Use Method: how to find how many cells make a battery. 8.2 ÷ 2.05 = 4.0, so four cells. Both electrodes gain mass, because solid $PbSO_4$ forms on each.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 5-6 of 8

5. Given a hydrogen fuel cell with the Data Booklet table rows $H_2 \rightarrow 2H^+ + 2e^-$ (0.00 V) and $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (+1.23 V). Write the overall reaction and find E°cell.

Overall: $2H_2 + O_2 \rightarrow 2H_2O$. E°cell = 1.23 − 0.00 = +1.23 V.

6. Given the electrolysis of NaCl(aq) in a pool system. Name the product at each electrode and show why $Cl_2 + 2OH^- \rightarrow Cl^- + OCl^- + H_2O$ is a disproportionation.

Hydrogen forms at the cathode and chlorine forms at the anode. In the reaction, Cl goes from 0 to −1 and from 0 to +1, so chlorine is both reduced and oxidized.

B13 · Worked practice

Batteries, fuel cells and chlorine for water | Questions 7-8 of 8

7. Given heating of copper(II) oxide with charcoal: $2CuO + C \rightarrow 2Cu + CO_2$. Name what is oxidized and reduced and the reducing agent.

Copper is reduced (+2 to 0). Carbon is oxidized (0 to +4). Carbon is the reducing agent.

8. Challenge: Given a hydrogen fuel cell with the anode half-reaction $H_2 \rightarrow 2H^+ + 2e^-$. The cell gives 2.00 A for 1.00 hour. Use the Faraday chain from B10. Find the mass of hydrogen gas used, in grams ($M(H_2) = 2.02$ g/mol). Use $F = 9.65 \times 10^4$ C/mol e^-.

Time: $t = 1.00\ h \times 3600\ s/h = 3600\ s$. Charge: $q = It = 2.00 \times 3600 = 7200$ C. Moles of electrons: $n(e^-) = 7200 \div (9.65 \times 10^4) = 0.074611$ mol. Ratio 2 e^- to 1 H_2: $n(H_2) = 0.074611 \div 2 = 0.037306$ mol. Mass: $m = 0.037306\ mol \times 2.02\ g/mol = 0.0754$ g. A fuel cell uses very little hydrogen by mass.