Quick facts

Redox in a galvanic cell occurs spontaneously, therefore the net cell reaction is exothermic. Chemical energy is converted to electrical energy.

Experiment 25 Date:

Aim: To determine the galvanic cell with the highest cell potential when comparing four specific metals.

Investigative question:

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Hypothesis:

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Apparatus:

- Four glass beakers
- Metal electrodes: Aℓ(s), Zn(s), Pb(s), Cu(s)
- Conducting wires, voltmeter
- ±125 cm^3 1 $mol \cdot dm^{-3}$ $A\ell(NO_3)_3(aq)$
- ±125 cm^3 1 $mol \cdot dm^{-3}$ $Cu(NO_3)_2(aq)$
- Filter paper
- Fine sanding paper or steel wool
- ±125 cm^3 1 $mol \cdot dm^{-3}$ $Zn(NO_3)_2(aq)$
- ±125 cm^3 1 $mol \cdot dm^{-3}$ $Pb(NO_3)_2(aq)$

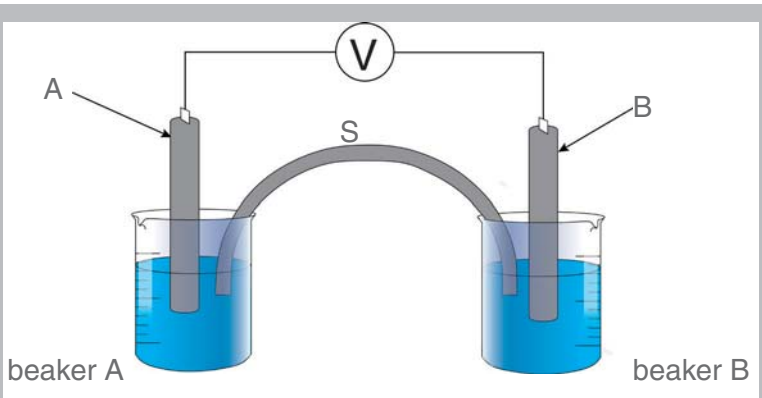
Safety measures:

- Prevent any skin or eye contact with the solutions. In the case of skin contact, rinse with excess water. In the case of eye contact, rinse well with clean water and acquire medical help.
- Ensure that the electrodes do not touch each other while the current flows through it.

Method:

1. Label the four glass beakers respectively as Zn, Aℓ, Pb and Cu.
2. Pour ±125 cm^3 1 $mol \cdot dm^{-3}$ $Zn(NO_3)_2(aq)$, ±125 cm^3 1 $mol \cdot dm^{-3}$ $A\ell(NO_3)_3(aq)$, ±125 cm^3 1 $mol \cdot dm^{-3}$ $Pb(NO_3)_2(aq)$ and ±125 cm^3 1 $mol \cdot dm^{-3}$ $Cu(NO_3)_2(aq)$ into the corresponding beakers, labelled A and B in the diagram.
3. Fold the filter paper (labelled S in the diagram) to form a strip of ± 1 cm wide. Dip the filter paper in the $KNO_3(aq)$.
4. Sand all the electrodes properly so that all deposits are removed and the metal looks shiny. Connect the Aℓ electrode (A), the Cu electrode (B) and voltmeter and place the electrodes in the corresponding $A\ell^{3+}(aq)$ and $Cu^{2+}(aq)$ solutions.

5. Place the filter paper that is soaked in the potassium nitrate solution, over the edges of the beakers as a salt bridge.
6. Immediately take the voltmeter reading.
7. Repeat steps 4 to 6 with Aℓ/Zn, Aℓ/Pb, Zn/Pb, Zn/Cu and Pb/Cu cells.



Observations:

Anode	Cathode	Cell potential (V)
$A\ell(s)/A\ell^{3+}(aq)$	$Cu(s)/Cu^{2+}(aq)$	
$A\ell(s)/A\ell^{3+}(aq)$	$Zn(s)/Zn^{2+}(aq)$	
$A\ell(s)/A\ell^{3+}(aq)$	$Pb(s)/Pb^{2+}(aq)$	
$Zn(s)/Zn^{2+}(aq)$	$Pb(s)/Pb^{2+}(aq)$	
$Zn(s)/Zn^{2+}(aq)$	$Cu(s)/Cu^{2+}(aq)$	
$Pb(s)/Pb^{2+}(aq)$	$Cu(s)/Cu^{2+}(aq)$	

Results:

Conclusions:

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2.2.3 Prediction of cell potential

Redox reactions between strong reducing and oxidising agents relative to one another occur spontaneously, while the redox reaction between weak reducing and oxidising agents relative to one another, will not occur spontaneously.

According to the position of these substances on the Table of Standard Electrode Potentials, it can be determined whether it is a strong or weak reducing or oxidising agent relative to one another.

The cell potential that was calculated in the previous section can also be used to predict the spontaneity of the redox reaction.

Step 1: Determine which substance acts as the reducing agent and read the electrode potential of the substance. The value then is E^0_{anode}.
Step 2: Determine which substance acts as the oxidising agent and read the electrode potential of the substance. The value then is $E^0_{cathode}$.
Step 3: Substitute these values into the equation $E^0_{cell} = E^0_{cathode} - E^0_{anode}$ to calculate the cell potential.
Step 4: If $E^0_{cell} > 0$, the reaction is spontaneous.
If $E^0_{cell} < 0$, the reaction is non-spontaneous.

Examples

Determine the cell potential for the following cell, then state whether the reaction is spontaneous or non-spontaneous:

$$Ag(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Ag^{+}(aq)$$

Step 1: According to the reaction equation the Ag(s) is oxidised, therefore Ag(s) is the anode.
Step 2: According to the reaction equation the $Cu^{2+}(aq)$ is reduced, therefore Cu(s) is the cathode.
Step 3: Use the Table of Standard Reduction Potentials to determine the values and to substitute them in the equation.

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$
$$= 0,34 - (0,80)$$
$$= -0,46 \text{ V}$$

Step 4: Decide on the basis of the E^0_{cell} value of the reaction if the reaction, as written, will occur spontaneously or not.
$\therefore E^0_{cell} < 0$
$\therefore$ the reaction is non-spontaneous.