Core practical 10: Construct electrochemical cells and measure electrode potentials

Objectives

- To construct an electrochemical cell
- To measure the electrode potential of a selection of electrochemical cells

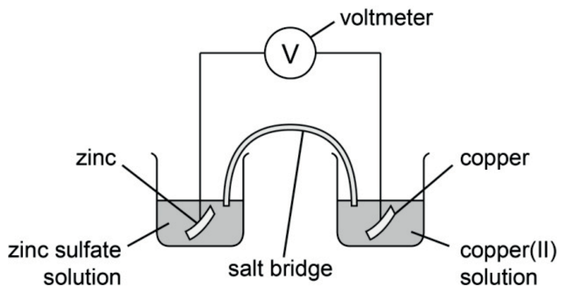
Safety

- Use eye protection.
- Zinc sulfate is harmful.
- 1.0 mol dm^{-3} iron(II) sulfate is harmful.
- Potassium nitrate is oxidising.

Specification links

- Practical techniques 4, 10, 11
- CPAC 1a, 2b, 3b, 4b

Diagram



Procedure

1. Clean the strips of zinc and copper using sandpaper.
2. Set up a zinc half-cell by pouring 50 cm^3 of zinc sulfate solution into the 100 cm^3 beaker and standing the strip of zinc in the beaker.
3. Set up a copper half-cell by pouring 50 cm^3 of copper(II) sulfate solution into a separate 100 cm^3 beaker and standing the strip of copper in the beaker.
4. Make an electrical connection between the two beakers by joining them with a strip of filter paper that has been dipped in a saturated solution of potassium nitrate (this acts as a salt bridge), as shown in the diagram.
5. Join the two metal strips with a voltmeter, using the connecting wires and crocodile clips.
6. Record the electrode potential of the [Zn(s) | Zn^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)] system. If the voltmeter gives a negative value, reverse the connections so that it gives a positive value.
7. Now, repeat steps 1–6 using the following combinations of metal/metal ion half-cells. Remember to clean the metal strips with sandpaper before use.

   [Zn(s) | Zn^{2+}(aq)] and [Fe^{2+}(aq) | Fe(s)]

   [Fe(s) | Fe^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)]

   [Zn(s) | Zn^{2+}(aq)] and [Ag^{+}(aq) | Ag(s)]

   [Cu(s) | Cu^{2+}(aq)] and [Ag^{+}(aq) | Ag(s)]

Notes on procedure

- Because the procedure is straightforward, students should be able to set it up without too many difficulties – however, they may need assistance in calculating $E^\ominus$ [Zn^{2+}(aq) | Zn(s)].

Answers to questions

1. $E^\ominus$ [Fe(s) | Fe^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)] = 0.78 V (or student's own result)

   $E^\ominus_{cell} = E^\ominus_{\text{right-hand half-cell}} - E^\ominus_{\text{left-hand half-cell}}$

   0.75 = 0.34 − $E^\ominus$ [Fe^{2+}(aq) | Fe(s)]

   The value for the iron half-cell is −0.44 V
2. The experiment was not carried out under standard conditions.
3. Silver nitrate is highly oxidising; an alternative answer is that silver nitrate is very expensive.
4. Magnesium reacts slowly with the water in the solution; raising the concentration of magnesium ions. The equilibrium will move to oppose this change and form more magnesium atoms.

Sample data

Theoretical values for these cells are:

[Zn(s) | Zn^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)] = 1.10 V

[Zn(s) | Zn^{2+}(aq)] and [Fe^{2+}(aq) | Fe(s)] = 0.32 V

[Fe(s) | Fe^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)] = 0.78 V

[Zn(s) | Zn^{2+}(aq)] and [Ag^{+}(aq) | Ag(s)] = 1.56 V*

[Cu(s) | Cu^{2+}(aq)] and [Ag^{+}(aq) | Ag(s)] = 0.46 V*

*This assumes that the solution of silver ions is 1.0 mol dm^{-3}.

Based on these theoretical values, the value for the zinc half-cell is $E^\ominus$ [Zn^{2+}(aq) | Zn(s)] = −0.76 V.

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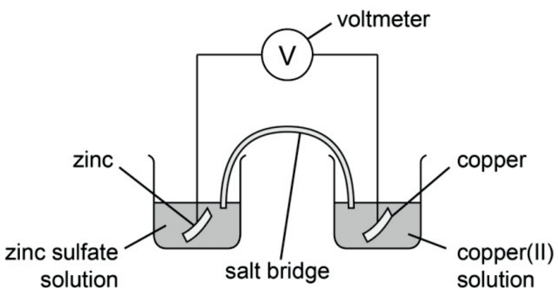
All the maths you need

- Substitute numerical values into algebraic equations using appropriate units for physical quantities:
  $E^{\ominus}_{\text{cell}} = E^{\ominus}_{\text{right-hand half-cell}} - E^{\ominus}_{\text{left-hand half-cell}}$

Equipment

- 50 cm^3 of 0.4 mol dm^{-3} zinc sulfate solution
- 50 cm^3 of 0.4 mol dm^{-3} copper(II) sulfate solution
- 50 cm^3 of 1.0 mol dm^{-3} iron(II) sulfate solution
- 50 cm^3 of 0.1 mol dm^{-3} silver nitrate solution
- saturated potassium nitrate solution
- distilled/deionised water
- one strip of each of the following: zinc; copper; iron and silver
- sandpaper
- four 100 cm^3 beakers
- strips of filter paper about 12 cm long
- 100 cm^3 measuring cylinder
- voltmeter (20 V) reading to 2 d.p.
- connecting wires and crocodile clips

Diagram



Procedure

1. Clean the strips of zinc and copper using sandpaper.
2. Set up a zinc half-cell by pouring 50 cm^3 of zinc sulfate solution into a 100 cm^3 beaker and standing the strip of zinc in the beaker.
3. Set up a copper half-cell by pouring 50 cm^3 of copper(II) sulfate solution into a separate 100 cm^3 beaker and standing the strip of copper in the beaker.
4. Make an electrical connection between the two beakers by joining them with a strip of filter paper that has been dipped in a saturated solution of potassium nitrate (this acts as a salt bridge), as shown in the diagram.
5. Join the two metal strips with a voltmeter, using the connecting wires and crocodile clips.
6. Record the electrode potential of the [Zn(s) | Zn^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)] system. If the voltmeter gives a negative value, reverse the connections so that it gives a positive value.
7. Now, repeat steps 1–6 using the following combinations of metal/metal ion half-cells. Remember to clean the metal strips with sandpaper before use.

   [Zn(s) | Zn^{2+}(aq)] and [Fe^{2+}(aq) | Fe(s)]

   [Fe(s) | Fe^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)]

   [Zn(s) | Zn^{2+}(aq)] and [Ag^{+}(aq) | Ag(s)]

   [Cu(s) | Cu^{2+}(aq)] and [Ag^{+}(aq) | Ag(s)]

Analysis of results

1. Record the electrode potential for each of the five cells you set up.
2. In the first cell you set up, the [Zn(s) | Zn^{2+}(aq)] and [Cu^{2+}(aq) | Cu(s)] system:

   $E^\ominus_{cell}$ = 1.10 V

   $E^\ominus$ [Cu^{2+}(aq) | Cu(s)] = +0.34 V
3. Use these values and the equation given in the 'All the maths you need' section to calculate $E^\ominus$ [Zn^{2+}(aq) | Zn(s)].

Learning tip

- To calculate $E^\ominus$ [Zn^{2+}(aq) | Zn(s)], substitute the values given into the equation:

  $E^\ominus_{cell} = E^\ominus_{\text{right-hand half-cell}} - E^\ominus_{\text{left-hand half-cell}}$

  $E^\ominus_{cell}$ = 1.10 V
- Note that the concentration of silver nitrate is 0.1 mol dm^{-3}, which is different to the other solutions. It is particularly dangerous to handle 1.0 mol dm^{-3} silver nitrate. The sample data shows the results for 1.0 mol dm^{-3} silver nitrate for you to compare.

Questions

1. Calculate $E^\ominus$ [Fe^{2+}(aq) | Fe(s)].
2. The electrode potential values for the cells you set up may be slightly different to theoretical values. Give a reason for this.
3. Give a reason why silver nitrate is not used as a 1.0 mol dm^{-3} solution.
4. [Mg^{2+}(aq) | Mg(s)] can also be used as a half-cell. Describe a problem that might be observed with this system.

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- Potassium nitrate is oxidising.
- Consult CLEAPSS Hazcards® 27C, 55B, 82, 87 and 108B. Perform a risk assessment using up-to-date information before this practical is carried out.

Equipment per student/group	Notes on equipment
50 cm^3 of 0.4 mol dm^{-3} zinc sulfate solution	Harmful Harmful to the environment
50 cm^3 of 0.4 mol dm^{-3} copper(II) sulfate solution	Low hazard at this concentration (harmful at higher concentrations).
50 cm^3 of 1.0 mol dm^{-3} iron(II) sulfate solution	Harmful
50 cm^3 of 0.1 mol dm^{-3} silver nitrate solution	Low hazard at this concentration (corrosive at higher concentrations). Harmful to the environment.
saturated potassium nitrate solution	Oxidising
distilled/deionised water	
one strip of each of the following: zinc; copper; iron and silver	
sandpaper	
four 100 cm^3 beakers	
strips of filter paper	About 12 cm long It would be helpful if these were already soaking in the saturated potassium nitrate solution at the start of the lesson.
100 cm^3 measuring cylinder	
voltmeter	To measure up to 20 V with an accuracy of 2 d.p.
connecting wires and crocodile clips	

Notes