ACTIVITY BRIEF

Cells and batteries

The science at work

Chemical reactions always absorb or release energy. The energy may be in the form of heat, light, electricity or sound. In some chemical reactions, there is a transfer of electrons between reactants to form products. By understanding how the movement of electrons can be controlled and harnessed, scientists can produce an electric current. Detailed investigations of this phenomenon enable scientists to develop cells and batteries that are efficient, safe and useful.

Your brief

You need carry out an investigation of cells and batteries. The aims are:

- to acquire and use scientific language associated with cells and batteries
- to find out about how electrochemical cells work and factors that can change their voltage output (potential difference between the positive and negative electrodes)
- to find out about commercial cells and batteries.

Task 1 Commercially available cells and batteries

Find out about commercially available cells and batteries, including:

- the different types that are on the market and what you need to consider when choosing a cell or battery for a particular purpose
- their care and maintenance
- why it’s important to recycle cells and batteries and how this is done.

You might find *Study sheet: Commercially available cells and batteries* useful in helping you organise this work. Collect information and make notes over a period of, say, 2-3 weeks. Then use this information to produce an illustrated leaflet that could be used to inform consumers about cells and batteries – what to buy for what purpose, how to look after them and how to dispose of them.

While working on this task, begin to make a glossary of terms used when talking about cells and batteries. Use *Study sheet: The language of cells and batteries*. You don’t need to complete it before starting work. Instead make a start and fill in answers as you work through the other parts of this activity.

Task 2 Making and testing electrochemical cells

Use metal/metal ion half-cells and a salt bridge to make an electrochemical cell. Measure the potential difference of some voltaic cells. Investigate factors that may or may not affect the magnitude of a cell's pd.

Use *Practical sheet: Making and testing electrochemical cells*.

Task 3 Emf and internal resistance

Find out about emf and internal resistance, and how they can be determined experimentally.

Use *Practical sheet: Emf and internal resistance*.

Task 4 Standard electrode potentials

Find out about emf, internal resistance and standard electrode potentials. Then calculate theoretical values for the emf of a cell and discuss why there is often a difference between calculated and experimental results.

Use *Study sheet: Standard electrode potentials*.

STUDY SHEET

Commercially available cells and batteries

You are going to produce an illustrated leaflet that could be used to inform consumers about cells and batteries – what to buy for what purpose, how to look after them and how to dispose of them.

To do this you must collect information about:

- the different types of cells and batteries on the market, how much they cost, what each type is best suited for, likely life expectancy
- their care and maintenance
- advantages/disadvantages of certain types – such as toxicity of cadmium, 'memory' etc.
- why it's important to recycle cells and batteries and how this is done.

You will probably want to find out about the following types of cells and batteries. (Remember: a battery is a collection of cells in series.)

Dry cell

General purpose disposable batteries, e.g.

- Zinc carbon
- Zinc chloride
- Alkaline manganese

Button cells, e.g.

- Silver oxide
- Zinc air
- Lithium manganese
- Mercuric oxide

Rechargeable, e.g.

- Nickel cadmium
- Nickel metal hydride
- Lithium ion

Wet cell

- Lead accumulator

There is a vast amount of information on the Internet, but you can also find out by just asking friends and relatives what they use and how they use them.

STUDY SHEET

The language of cells and batteries

Term	**Definition or explanation**
cell	
battery	
positive electrode	
negative electrode	
electrode potential	
half-cell	
electrochemical cell	
salt bridge	
half-ion reactions	
potential	
potential difference	
(conventional) current	
emf	
terminal potential difference	
internal resistance	
standard hydrogen electrode	

PRACTICAL SHEET

Making and testing electrochemical cells

IMPORTANT

Before starting an experiment make sure you:

- know the purpose of the experiment and are organised
- have tables ready to record observations and measurements
- note measurements accurately, together with the degree of accuracy of the apparatus and instruments used
- note any conditions and materials that may affect results
- identify hazards and prepare a risk assessment.

You can learn about how cells and batteries produce electricity and factors that affect their performance by studying some simple electrochemical cells.

A Daniell cell

A Daniell cell is an **electrochemical cell** consisting of a zinc electrode in zinc sulfate solution and a copper electrode in copper sulfate solution, linked by a salt bridge. Use the method below to make one.

Equipment and materials

- a piece of zinc sheet or rod
- a piece of copper sheet or rod

1 mol dm^{-3} copper sulfate solution [HARMFUL]

1 mol dm^{-3} zinc sulfate solution [ZINC SULFATE IS AN IRRITANT]

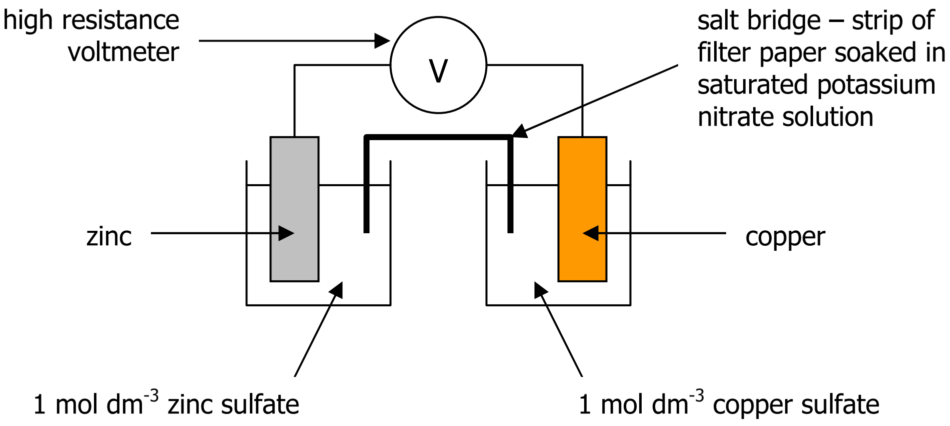
- 2 x sample tubes or small beakers
- high resistance voltmeter
- connecting wire and crocodile clips (or similar)

strip of filter paper soaked in saturated potassium nitrate solution [SOLID POTASSIUM NITRATE IS AN OXIDANT]

Health and safety

A risk assessment must have been carried out.

Method



a Set up the apparatus shown on the previous page.

b Connect the voltmeter leads to the two electrodes. Read the voltmeter. Disconnect and connect the other way round. Again, read the voltmeter.

Which electrode is the negative electrode and which is the positive electrode?

c With the voltmeter connected correctly (with the voltage positive), remove the salt bridge. Note what happens.

d Replace the salt bridge and leave the cell to run for about 20 minutes. Record, in a suitable table, the voltmeter reading (**potential difference**) every two minutes and your observations.

Note: you may need to moisten the salt bridge with saturated potassium nitrate solution if it starts to dry out.

Questions

1 If zinc is put into a solution of copper sulfate, the zinc slowly dissolves and the solution becomes a lighter blue. (You might want to try this.) The reaction mixture also gets hot because the reaction is exothermic, showing that energy is released. Here is the equation for the reaction:

$$Zn(s) + CuSO_4(aq) \rightleftharpoons ZnSO_4(aq) + Cu(s)$$

or written just in terms of the ions involved

$$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$$

a Explain why this is a redox reaction.

b Complete theses two half-ion equations and show how they can be ‘added’ together to give the full ion equation:

$Zn(s) \rightleftharpoons$ ________________

________________ $\rightleftharpoons Cu(s)$

Full ion equation:

____

2 The cell may be described as being made up of two **half cells**. Each has a metal (the electrode) in a solution of its ions. In this cell, zinc is the negative electrode and copper is the positive electrode.

Describe, using ionic equations, what happens in the cell. Hint: think about the equations you wrote in question 1.

Explain how your observations support this.

3. Draw the electrochemical cell and circuit. Add the labels 'negative terminal' and 'positive terminal'.

   Show the direction of:

   a. the flow of electrons
   b. the (conventional) current.

4. Explain the purpose of the **salt bridge** and how it works.

More electrochemical cells

Set up a number of electrochemical cells by pairing up various metal/metal ion half cells using these metals in 1 mol dm^{-3} solutions of their salts:

magnesium zinc copper tin lead nickel

Health and safety note: some zinc compounds are HAZARDOUS, some copper compounds are HARMFUL, lead and some of its compounds are TOXIC even in very dilute solutions, nickel and most of its compounds are HARMFUL even in dilute solutions they can cause skin sensitivity.

For each cell you make, record the initial potential difference and monitor this over a period of time (as you did with the Daniell cell).

Questions

1. List the cells you constructed in order of increasing potential difference. Explain the sequence by referring to the reactivity of the metals.
2. Deduce a general rule relating the potential difference of a cell to the position of the metals in the electrochemical series.
3. Write half-ion equations and an overall ionic equation for each electrochemical cell.

Batteries

How can we get higher voltages? The answer is to put a number of cells in series. In other words, make a **battery**.

Make four Daniell cells.

Measure and record the voltage of one cell (as you did earlier). Now put four cells in series. Using a voltmeter with a flying lead, measure the potential difference across:

- one cell
- two cells
- three cells
- four cells.

Record you findings.

Questions

1. Describe the difference in voltage output from the various numbers of cells in series.
2. What voltage would you expect from two cells set up in parallel? Test your prediction.

Cell performance

You will have seen that the voltage of an electrochemical cell decreases as the chemicals get used up. In this next part you will investigate factors that might influence the potential difference of an electrochemical cell.

The cell you will use consists of the two half cells: $Mg(s)/Mg^{2+}(aq)$ and $Cu(s)/Cu^{2+}(aq)$

Make an electrochemical cell using cleaned magnesium ribbon, copper foil, 1 mol dm^{-3} magnesium sulfate and 1 mol dm^{-3} copper sulfate. The other equipment you need is the same as for making a Daniell cell.

Measure the potential difference of the cell at room temperature. Record both the temperature and the pd.

Predictions

Now say how you think the potential difference might change if:

A the size and shape of the electrodes are changed

B the concentrations of the salts are changed:

- the concentrations of $MgSO_4(aq)$ and $CuSO_4(aq)$ are equal but less than 1 mol dm^{-3}
- the concentrations of $MgSO_4(aq)$ and $CuSO_4(aq)$ are equal but greater than 1 mol dm^{-3}
- the concentration of $MgSO_4(aq)$ is less than the concentration of $CuSO_4(aq)$
- the concentration of $MgSO_4(aq)$ is greater than the concentration of $CuSO_4(aq)$

C for each of the situations described in **B**, the temperature is:

- less than room temperature
- greater than room temperature.

Testing the predictions

Plan a series of experiments to test your predictions. Carry out a risk assessment for the method you plan to use. Check your plan and risk assessment with your teacher.

And now some theory

The Nernst equation describes how potential difference varies with conditions:

$$E_{cell} = E^o_{cell} - \frac{RT}{nF}\ln Q$$

(If you are not sure what 'ln' means, ask you teacher)

where E_{cell} = cell potential difference / V

E^o_{cell} = cell potential difference under standard conditions (298 K and 100 kPa) / V

R = gas constant (8.314 J K^{-1} mol^{-1})

T = temperature / K

n = number of electrons transferred

F = Faraday's constant (96 485.3415 C)

Q = thermodynamic reaction quotient

You don't need to know how this equation was derived or the precise meaning of the terms.

Q may be written as:

concentration of metal ions in negative half cell ($c_{negative}$) ÷ concentration of metal ions in positive half cell ($c_{positive}$)

So the Nernst equation may be rewritten as:

$$E_{cell} = E^{o}_{cell} - \frac{RT}{nF} \ln \frac{c_{negative}}{c_{posiitve}}$$

In a Daniell cell, $c_{negative}$ = concentration of zinc ions and $c_{positive}$ = concentration of copper ions

Questions

1 In the magnesium/copper cell you made, which ions are in the negative half cell and which are in the positive half cell?

2 Comment on the results obtained from the experiments designed to test your predictions in relation to the Nernst equation.

3 Use the Nernst equation to calculate the potential difference when:

   a $[Mg^{2+}(aq)] = 0.1$ mol dm^{-3}; $[Cu^{2+}(aq)] = 0.1$ mol dm^{-3}; temperature = 25 ^{o}C
   b $[Mg^{2+}(aq)] = 1$ mol dm^{-3}; $[Cu^{2+}(aq)] = 1$ mol dm^{-3}; temperature = 5 ^{o}C
   c $[Mg^{2+}(aq)] = 1$ mol dm^{-3}; $[Cu^{2+}(aq)] = 1$ mol dm^{-3}; temperature = 95 ^{o}C
   d $[Mg^{2+}(aq)] = 1$ mol dm^{-3}; $[Cu^{2+}(aq)] = 0.01$ mol dm^{-3}; temperature = 25 ^{o}C
   e $[Mg^{2+}(aq)] = 1$ mol dm^{-3}; $[Cu^{2+}(aq)] = 0.01$ mol dm^{-3}; temperature = 5 ^{o}C
   f $[Mg^{2+}(aq)] = 1$ mol dm^{-3}; $[Cu^{2+}(aq)] = 0.01$ mol dm^{-3}; temperature = 95 ^{o}C
   g $[Mg^{2+}(aq)] = 0.01$ mol dm^{-3}; $[Cu^{2+}(aq)] = 1$ mol dm^{-3}; temperature = 25 ^{o}C
   h $[Mg^{2+}(aq)] = 0.01$ mol dm^{-3}; $[Cu^{2+}(aq)] = 1$ mol dm^{-3}; temperature = 5 ^{o}C
   i $[Mg^{2+}(aq)] = 0.01$ mol dm^{-3}; $[Cu^{2+}(aq)] = 1$ mol dm^{-3}; temperature = 95 ^{o}C
   j $[Mg^{2+}(aq)] = 1$ mol dm^{-3}; $[Cu^{2+}(aq)] = 0.0001$ mol dm^{-3}; temperature = 25 ^{o}C
   k $[Mg^{2+}(aq)] = 0.0001$ mol dm^{-3}; $[Cu^{2+}(aq)] = 1$ mol dm^{-3}; temperature = 25 ^{o}C

PRACTICAL SHEET

Emf and internal resistance

To evaluate the usefulness of any cell as a source of electricity, two properties need to be known: its electromotive force (emf) and its internal resistance.

An electrochemical cell converts chemical energy into electrical energy. The cell's **emf** (**electromotive force**) is the maximum energy a cell can give to each coulomb of charge when it flows through a cell. (1 volt = 1 joule per coulomb)

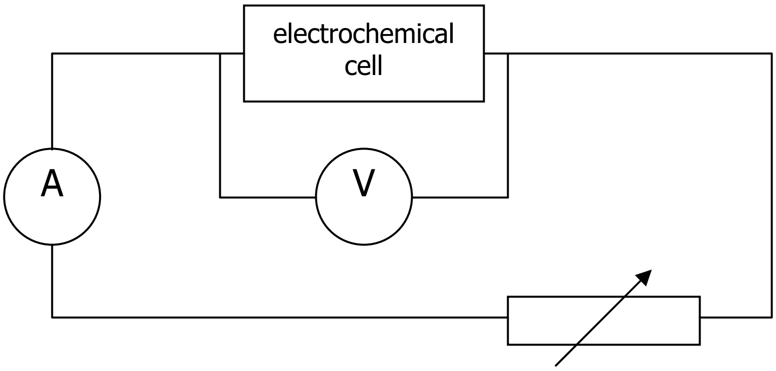
An electrochemical cell is made of various materials. These have an electrical resistance. Together they make up a cell's **internal resistance,** $\boldsymbol{r}$. This causes a voltage to develop across the cell when a current flows. This 'lost' voltage is not available to the rest of the circuit. So ideally a cell has a low internal resistance.

The cell's **terminal potential difference** (voltage across the cell's terminals), $\boldsymbol{V_{t.p.d.}}$, is given by:

$V_{t.p.d.}$ = emf - Ir

where I = current (A)

Now look at this circuit:



1 Explain how it might be used to enable a graph of terminal potential difference (V) against current (A) to be plotted.

2 Explain how this graph may be used to determine the emf and internal resistance of the cell.

Now try the experiment for real

- Set up the circuit shown above. Use a C cell for the electrochemical cell. You will also need an ammeter and voltmeter (or two multimeters), a rheostat (approx. 10 Ω) and leads.
- Set the rheostat to its maximum resistance.
- Record the terminal potential voltage and the current.

- Reduce the rheostat to its lowest resistance (zero) in ten equal steps. Each time, record terminal potential voltage and the current.

  Note: Do not leave the circuit connected for long when the current is high. This will run down the cell quickly.
- Plot a graph of terminal potential voltage (V) against current (A).
- Use your graph to determine the emf of the C cell and its internal resistance.

STUDY SHEET

Standard electrode potentials

You've seen that electrochemical cells consist of two half cells.

Each half cell has a **potential**. This is usually called the **electrode potential**. It's a measure of how easily metal atoms shed their electrons to form hydrated ions. Each metal/metal ion half cell has its own **electrode potential** and the more reactive the metal, the higher this potential.

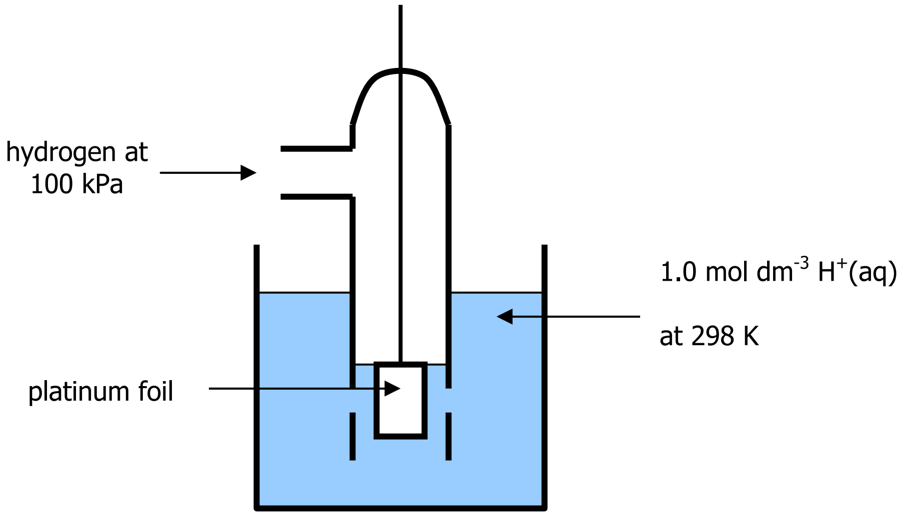
This potential can't be measured. What we can measure is the voltage (**potential difference**) of a cell made from the two half cells. You did this if you used *Practical sheet: Making and testing electrochemical cells*.

We can't measure potential so we need to measure everything else relative to a standard potential.

$$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$$

As potential differences depend on conditions (which you will have seen, again, in *Practical sheet: Making and testing electrochemical cells*), the first thing we need to do is specify our standard conditions. These are: H^+ ion concentration 1.0 mol dm^{-3}; 298 K (25 ^oC); 100 kPa (normal atmospheric pressure).

Below is the set up for the half cell we call a **standard hydrogen electrode**. It is assigned a standard electrode potential of zero (0).



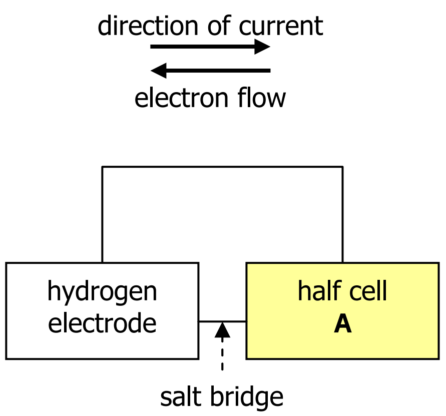
Now we compare all other half cells with this. This is how …

- Make an electrochemical cell by connecting a metal/metal ion half cell (under standard conditions) to a standard hydrogen electrode. Measure the cell emf.
- From the direction of the current, work out which is the negative electrode and which is the positive electrode in the cell. Remember: by convention, current flows in the opposite direction to the flow of electrons.

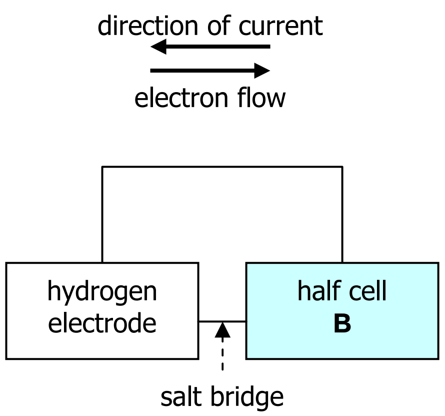
Questions

1 Look at these diagrams:

Cell 1



Cell 2



Complete:

Cell 1

negative electrode is _______________ and positive electrode is _________________

Cell 2

negative electrode is _______________ and positive electrode is _________________

2 The reading on the voltmeter is the emf of the half cell since the hydrogen half cell has zero potential. Complete these sentences:

a If the half cell is the **negative electrode** (electrons flow ______ the hydrogen electrode), the half cell standard potential is given a _______ sign.

b If the half cell is the **positive electrode** (electrons flow ______ the hydrogen electrode), the half cell standard potential is given a ______ sign.

3 The standard potentials are usually given as the potential of the reduction reaction. Here are some examples:

Ionic half equation	**Standard potential (V)**
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	-2.37
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	-1.66
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	-0.76
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

In what direction do electrons flow when

- a Mg(s)/Mg^{2+}(aq) half cell is connected to a Ag(s)/Ag^{+}(aq) half cell
- a Cu(s)/Cu^{2+}(aq) half cell is connected to an Al(s)/Al^{3+}(aq) half cell
- an Al(s)/Al^{3+}(aq) half cell is connected to a Mg(s)/Mg^{2+}(aq) half cell

4 Finally, if we know the standard potentials of two half cells, we can calculate the standard potential difference for the cell made from them (i.e. its emf).

cell standard potential difference =
positive electrode standard potential – negative electrode standard potential

For example:

For a cell made from a Mg(s)/Mg^{2+}(aq) half cell and a Cu(s)/Cu^{2+}(aq) half cell

Mg(s)/Mg^{2+}(aq) is the negative electrode half cell with standard potential -2.37 V

Cu(s)/Cu^{2+}(aq) is the positive electrode half cell with standard potential +0.34 V

So the standard potential of the cell = +0.34 – (-2.37) = 2.71 V

Calculate the standard potential of cells consisting of:

- Mg(s)/Mg^{2+}(aq) and Ag(s)/Ag^{+}(aq) half cells
- Cu(s)/Cu^{2+}(aq) and Ag(s)/Ag^{+}(aq) half cells
- Al(s)/Al^{3+}(aq) and Mg(s)/Mg^{2+}(aq) half cells

Teacher notes

From the assessment evidence grid for OCR Unit 12 Electrons in action:

***You need to produce** evidence of your investigation into the principles and applications of electrochemical changes [**50** marks].*

*This evidence needs to include**:***

***AO1:** a presentation outlining the applications of stated electrochemical changes [**10**]*

***AO2:** a comparison of commercial cells: non-rechargeable, rechargeable and fuel, including construction, resources, uses, sustainability, efficiency, safety and environmental issues [**14**]*

***AO3:** practical investigations into:*

(a) the factors which can change the potential difference of a cell and those which have no effect

*(b) the factors which affect the efficiency of a simple laboratory experiment in which an object is copper plated [**25**].*

This activity will allow students to provide evidence for AO2 and AO3(a).

For AO1, students gain the highest marks if they

- *demonstrate research into the production of electricity, using the full range of examples given, selecting and interpreting information and presenting it clearly*

For AO2, students gain the highest marks if they

- *describe a wide range of cells, make all comparisons, give a full explanation and interpretation of information*
- *carry out complex calculations of EMF of cells, quantities of charge and mass of products*
- *obtain and use data to compare the efficiency of commercial cells and obtain correct solutions to the appropriate degree of accuracy*

For AO3(a), students gain the highest marks if they

- *produce their own detailed risk assessments, consistent with COSHH guidelines*
- *carry out a wide range of measurements of EMF of cells*
- *consider and change a range of conditions to obtain corresponding sets of results for the measurement of EMF of cells – at least **one** set of results show no effect*
- *explain any practical techniques that will improve results*
- *work with an appropriate degree of accuracy*
- *make and record relevant observations and measurements from the above experiments, using precision in your measurements*
- *display the data obtained accurately in a range of ways*
- *interpret the results in detail and draw conclusions*
- *evaluate their procedures and suggest alternatives.*

Please note: To meet CLEAPSS Health and Safety standards, when writing risk assessments students should refer to the *Management of Health and Safety at work Regulations* as well as COSHH guidelines.

Overall teaching strategy

Task 1 Commercial cells and batteries

Students might be given this at the beginning of the unit, with instructions to collect information and make notes during the time allocated to the unit, culminating with the production of the information leaflet in the last 2-3 weeks.

It's suggested that each student is given a copy of *Study sheet: The language of cells and batteries*. They might be asked to fill in the boxes where they already know the answers and complete the others during the course of the unit. They should also check the ones they think they already know.

The remaining tasks (2-4) might run sequentially. There is plenty of practical work.

Making and testing electrochemical cells

One assessment requirement for the unit is that students must carry out:

> *Practical investigations into: (a) the factors which can change the potential difference of a cell and those which have no effect*

with the highest marks gained if they show that they have

- *produced their own detailed risk assessments, consistent with COSHH guidelines*
- *carried out a wide range of measurements of EMF*
- *considered and changed a range of conditions to obtain corresponding sets of results for measurement of EMF – at least* ***one*** *set of results show no effect*
- *explained any practical techniques that will improve results*
- *worked with an appropriate degree of accuracy*
- *made and recorded relevant observations and measurements from the above experiments, using precision in your measurements*
- *displayed the data obtained accurately in a range of ways.*
- *interpreted the results in detail and draw conclusions*
- *evaluated their procedures and suggest alternatives.*

To gain the highest marks it is clear that students need to display quite a bit of autonomy.

Before students start they should know about dynamic equilibrium and redox reactions (in terms of electron transfer).

They can learn about how cells and batteries produce electricity and factors that affect their performance by studying some simple electrochemical cells.

A Daniell cell

As an alternative to using half cells linked with a salt bridge, a zinc rod, together with the zinc sulfate solution, could be put in Visking tubing, which is then placed in a beaker containing copper sulfate solution and the copper electrode.

The initial voltmeter reading should be about 1.1 V. The voltage should decrease slowly as the zinc ion concentration increases and the copper ion concentration decreases (see Nernst equation later). Zinc dissolves and copper sulfate solution becomes lighter blue.

Answers

1

- a Electrons are transferred from zinc atoms to copper ions. Zinc is oxidised (loses electrons) and copper ions are reduced (gain electrons)
- b Complete theses two half-ion equations and show how they can be 'added' together to give the full ion equation:

  $Zn(s) \rightleftharpoons Zn^{2+}(aq) + 2e^-$

  $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$

  Full ion equation:

  $Zn(s) + Cu^{2+}(aq) + \cancel{2e^-} \rightarrow Zn^{2+}(aq) + \cancel{2e^-} + Cu(s)$

  The two moles of electrons of either side of the equation cancel out leaving:

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

2 $Zn/Zn^{2+}(aq)$ half cell: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

$Cu/Cu^{2+}(aq)$ half cell: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

So electrons flow from the zinc electrode (the negative terminal) to the copper electrode (the positive terminal).

Zinc dissolves and copper sulfate solution becomes lighter blue as more copper is deposited.

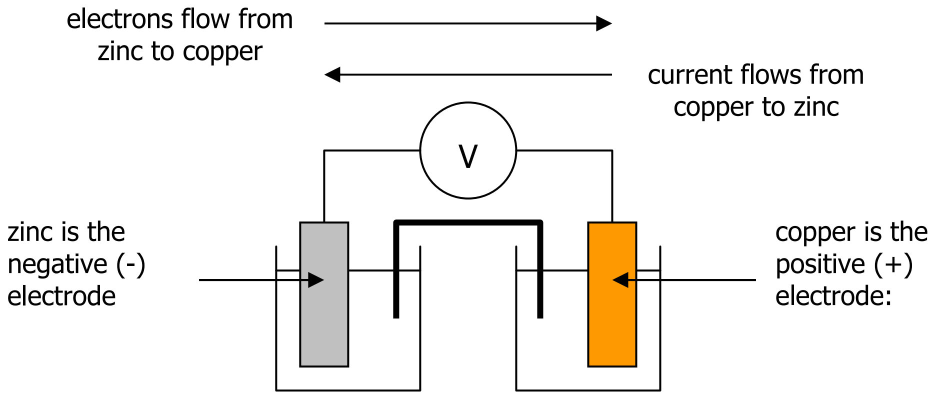
3

- a the flow of electrons is from the zinc electrode to the copper electrode
- b the direction of the current is from the copper electrode to the zinc electrode

Conventional current assumes that current flows out of the positive terminal, through the circuit and into the negative terminal of the source. This was the convention chosen during the discovery of electricity.

Electron flow is what actually happens and electrons flow out of the negative terminal, through the circuit and into the positive terminal of the source.

Both terms are used by industry.



4 It makes an electrically conducting connection between the solutions in the half cells.

More electrochemical cells

Answers

1 The bigger the reactivity difference, the higher the pd.

Batteries

Suggest that students work in groups of four. Each student makes a Daniell cell.

Answers

1 The potential difference is the sum of the pd's of the individual cells that are in series.

2 Same as for one cell – but the pair contain more chemicals to react and so can produce electricity for longer.

Cell performance

Changing the dimensions of the electrodes has no effect.

The effects of changing concentration and temperature are very small, so a good quality voltmeter will be needed.

When testing their predictions, students might find it odd that changing temperature only has an effect when the concentrations of ions in the two half cells are unequal. The explanation can be discussed after they have met the Nernst equation in the next section. (With equal concentrations, the ratio = 1, and $\ln 1 = 0$.)

And now some theory

You might not want students to see this until after they have made their predictions.

From the Nernst equation we can make predictions (see answers to question 3 below). For example, at 25 °C:

$[Mg^{2+}(aq)]$ / mol dm^{-3}	1	0.1	0.01	0.001	1	1	1
$[Cu^{2+}(aq)]$ / mol dm^{-3}	1	1	1	1	0.1	0.001	0.001
pd / V	2.71	2.740	2.769	2.799	2.680	2.651	2.621

Answers

1 Mg^{2+} and Cu^{2+} respectively

2 Appropriate comments

3 Use the Nernst equation to calculate the potential difference when:

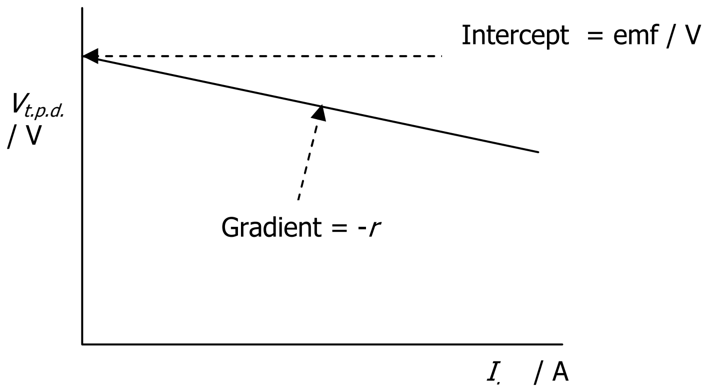
  a 2.71 V
  b 2.71 V
  c 2.71 V
  d 2.651 V
  e 2.655 V
  f 2.637 V
  g 2.769 V
  h 2.765 V
  i 2.783 V
  j 2.592 V
  k 2.828 V

Students could use Excel to calculate E_{cell} for different cells under different conditions.

Emf and internal resistance

$V_{t.p.d.}$ = emf - Ir

Plot a graph of the cell's terminal potential difference (*y*-axis) against the current flowing in the circuit (*x*-axis). The intercept on the *y*-axis gives the cell's emf and the gradient of the straight line is the negative value for the cell's internal resistance.



Standard electrode potentials

The relationship between the content described in the specification and the assessment requirements is unclear in places. For example, the content says

- *explain the term 'standard hydrogen electrode' and describe how it is measured*

While this is what appears in the specification it does not make sense. One cannot measure an electrode, nor is it possible to measure the hydrogen electrode potential. Electrode potentials are determined by measuring them against a hydrogen electrode which is assigned a value of 0 volts.

- *use standard electrode potentials to calculate EMF of cells.*

However, the assessment requirements include

AO1: *a presentation outlining the applications of stated electrochemical changes*

AO2: *a comparison of commercial cells: non-rechargeable, rechargeable and fuel, including construction, resources, uses, sustainability, efficiency, safety and environmental issues*

It isn't clear where standard electrode potentials fit. One of the marking criteria for AO2 does say *you will carry out complex calculations of EMF of cells*, but this is ambiguous. It could mean using the Nernst equation (see *Practical sheet: Making and testing electrochemical cells*).

Students may be confused about the difference between emf and potential difference. *Practical sheet: emf and internal resistance* may help. They need to know that emf can be measured by finding the potential difference across the terminals when no current is flowing.

Answers

1 Cell 1

negative electrode is **A** and positive electrode is the hydrogen electrode

Cell 2

negative electrode is the hydrogen electrode and positive electrode is **B**

2 Complete these sentences (missing words are <u>underlined</u>):

If the half cell is the **negative electrode** (electrons flow <u>to</u> the hydrogen electrode), the half cell standard potential is given a <u>negative</u> sign.

If the half cell is the **positive electrode** (electrons flow <u>from</u> the hydrogen electrode), the half cell standard potential is given a <u>positive</u> sign.

3

- a $Mg(s)/Mg^{2+}(aq)$ half cell is connected to a $Ag(s)/Ag^{+}(aq)$ half cell

  Answer Mg to Ag
- a $Cu(s)/Cu^{2+}(aq)$ half cell is connected to an $Al(s)/Al^{3+}(aq)$ half cell

  Answer Al to Cu
- an $Al(s)/Al^{3+}(aq)$ half cell is connected to a $Mg(s)/Mg^{2+}(aq)$ half cell

  Answer Mg to Al (Mg is the more –ve)

4 Finally, if we know the standard potentials of two half cells, we can calculate the emf of a cell made from them.

– $Mg(s)/Mg^{2+}(aq)$ and $Ag(s)/Ag^{+}(aq)$ half cells	3.17 V	
– $Cu(s)/Cu^{2+}(aq)$ and $Ag(s)/Ag^{+}(aq)$ half cells	0.46 V	[0.80 – 0.34]
– $Al(s)/Al^{3+}(aq)$ and $Mg(s)/Mg^{2+}(aq)$ half cells	0.71 V	[-1.66 – (-2.37)]