

Practical Booklet 7

Cell potentials and the Nernst equation

Cambridge International AS & A Level Chemistry 9701

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Introduction

Practical work is an essential part of science. Scientists use evidence gained from prior observations and experiments to build models and theories. Their predictions are tested with practical work to check that they are consistent with the behaviour of the real world. Learners who are well trained and experienced in practical skills will be more confident in their own abilities. The skills developed through practical work provide a good foundation for those wishing to pursue science further, as well as for those entering employment or a non-science career.

The science syllabuses address practical skills that contribute to the overall understanding of scientific methodology. Learners should be able to:

- plan experiments and investigations
- collect, record and present observations, measurements and estimates
- analyse and interpret data to reach conclusions
- evaluate methods and quality of data, and suggest improvements.

The practical skills established at AS Level are extended further in the full A Level. Learners will need to have practised basic skills from the AS Level experiments before using these skills to tackle the more demanding A Level exercises. Although A Level practical skills are assessed by a timetabled written paper, the best preparation for this paper is through extensive hands-on experience in the laboratory.

The example experiments suggested here can form the basis of a well-structured scheme of practical work for the teaching of AS and A Level science. The experiments have been carefully selected to reinforce theory and to develop learners' practical skills. The syllabus, scheme of work and past papers also provide a useful guide to the type of practical skills that learners might be expected to develop further. About 20% of teaching time should be allocated to practical work (not including the time spent observing teacher demonstrations), so this set of experiments provides only the starting point for a much more extensive scheme of practical work.

Guidance for teachers

Aim

To investigate the magnitude of cell potentials using a variety of half cells, and to investigate the effect of change of concentration on the cell potential.

Outcomes

Syllabus section 6.1, 6.3(c), (d), (e), (h) and (i) as well as experimental skills 1, 2, 3, and 4.

Further work: syllabus section 6.2(b)

Syllabus sections 12.1 (h) and 12.2 (m)

Skills included in the practical

A Level skills	How learners develop the skills
Planning	identify a safe and efficient procedure that when followed would lead to a reliable result describe the steps necessary to carry out the procedure suggest appropriate volumes and concentrations describe how to vary the independent variable and how the dependent variable is to be measured
Analysis	calculate concentrations of solutions used in the half cells use the Nernst equation describe the patterns and trends in their data
Evaluation	identify sources of error in the procedure suggest ways to improve the accuracy of the procedure suggest ways in which to extend the investigation to answer a new question
Conclusions	determine electron flow in the external circuit deduce whether the data collected supports the Nernst equation make scientific explanations of the data prediction of relative strengths of oxidising agents (and of reducing agents)

This practical provides an opportunity to build on essential skills introduced at AS Level.

AS Level skills	How learners develop the skills
MMO collection	set up and use apparatus to an appropriate level of precision
PDO recording	record the pd, half-cell description and sign with appropriate headings and units

Methods A and B

- **Learners must wear eye protection for these investigations.**
- The simplest type of electrode consists of a metal dipping into an aqueous solution of its ions. The metal and ions form an equilibrium: $M(s) \rightleftharpoons M^{2+}(aq) + 2e^-$ (for metals that form 2+ ions).

If the metal reacts, electrons are produced on the surface of the metal. The more reactive the metal is, the more electrons are produced, so the more negative the metal and the electrode become.
- A simple cell is made by connecting together two metal electrodes (half-cells), using a salt bridge. This makes electrical contact between the two electrodes, with no mixing of the solutions.
- The voltage generated by the cell depends mainly on the difference in reactivity of the two metals. The concentration of the solution of the metal ions into which the metals dip also affects the voltage slightly.
- The equilibrium between a metal and its ions is a redox reaction, so other redox pairs can be used in a half-cell (eg $Fe^{2+}(aq) / Fe^{3+}(aq)$). In this case an inert metal, platinum, is used to provide the pathway for electrons in and out of the half-cell.
- When carrying out the experiments, learners should be aware that a fresh salt bridge is needed for each pair of half-cells. They should also take care that crocodile clips (or other connectors) do not dip into the solutions of ions under test. They should check the connections for corrosion.
- In method B, learners will plan an experiment to investigate changing the concentration of one of the solutions in a cell or half-cell, and determine the effect this has on the pd of the cell formed. In this way they will investigate the Nernst equation. (It may be necessary to prompt learners to alter concentrations by factors of 10 as the Nernst equation includes a \log_{10} term.)

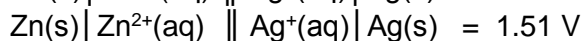
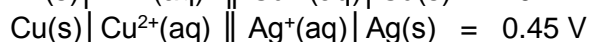
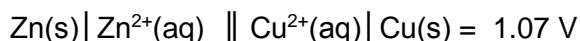
Results

- Learners should tabulate results clearly, showing which is the positive half of the cell and recording the value of the cell pd / V.
- All voltmeter readings should be shown to the same number of decimal places.

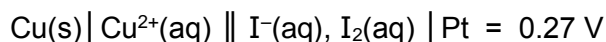
Interpretation and evaluation

- Discussion can take place to determine which term is more appropriate: cell emf or cell pd.
- The reasons for differences between the experimentally determined cell potentials and the E^\ominus values can be discussed (concentrations not 1.00 mol dm^{-3} , temperature not 298 K, some current flow).
- The methods to reduce the concentrations accurately can be discussed. A simple method is to use measuring cylinders or more precisely calibrated apparatus can be used instead.
- The effect of changes in concentration on the $M^{n+}(aq) + ne^- \rightleftharpoons M(s)$ equilibrium can be discussed.

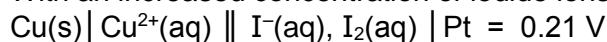
Typical results



A simple demonstration of the effect of changing one concentration in a half-cell of two solutions is to use the $\text{Pt} | \text{I}_2(\text{aq}), \text{I}^{-}(\text{aq})$ half-cell connected to one of the others and take a reading. Then add solid potassium iodide to the half-cell, stir and take a new voltmeter reading. (The concentration of aqueous iodine was not known.)



With an increased concentration of iodide ions,



Calculation

If the Cu^{2+} solution is diluted by a factor of 10, the cell potential should decrease (become less positive) by approximately 0.03 V. Hence the voltmeter reading with the original Zn half cell and the new Cu half-cell should be 1.04 V from the results above.

Further work

The connection between cells producing electrical energy from chemical potential energy and electrolysis where electrical energy is used to produce chemical energy can be discussed (reversible reactions). The ability of metals to donate electrons and the ease of reduction of their ions can be linked. Similarly, the oxidising ability of non-metal systems such as the halogens and the ease of producing the non-metal by electrolysis can be linked.

Information for technicians

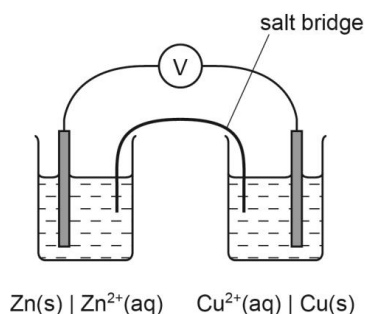
Each learner will require:

- (a) Eye protection
- (b) 1 x 250 cm³ beaker
- (c) 6 x 100 cm³ beaker (or smaller)
- (d) 1 x 25 cm³ measuring cylinder
- (e) 1 x high resistance d.c. voltmeter or multimeter (reading to **at least** 0.1 V)
- (f) 2 x crocodile clips
- (g) 2 x connecting wires
- (h) filter paper, scissors and forceps
- [O] (i) 25 cm³ saturated aqueous potassium nitrate(V)
- [MH] [N] (j) 50.0 cm³ 1.0 mol dm⁻³ copper(II) sulfate
- [MH] [N] (k) 50.0 cm³ 1.0 mol dm⁻³ zinc sulfate
- [C] [N] [HH] (l) 50.0 cm³ 1.0 mol dm⁻³ magnesium sulfate
- [MH] (m) 50.0 cm³ 1.0 mol dm⁻³ iron(II) sulfate
- [C] [MH] (n) 50.0 cm³ 1.0 mol dm⁻³ iron(III) chloride
- [C] [O] (o) 50.0 cm³ 0.05 mol dm⁻³ silver nitrate
- [N] (p) 50.0 cm³ 0.02 mol dm⁻³ iodine dissolved in approx 0.05 mol dm⁻³ potassium iodide
- (q) 1 x copper strip (electrode) with surface cleaned with fine sandpaper
- [F] [N] (r) 1 x zinc (electrode) with surface cleaned with fine sandpaper
- [F] (s) 1 x magnesium ribbon (electrode) with surface cleaned with fine sandpaper
- (t) access to silver electrode or silver wire (a silver ring may be used)
- (u) access to platinum electrode
- (v) potassium iodide

Additional Instructions








The solutions may be retained and reused provided they have not been mixed or otherwise contaminated and there has been no significant current flow. (The amount of KNO₃ entering the solutions is minimal.)

Equipment set-up



Technician's notes, *continued*

Hazard symbols

 <p>GHS02 (<i>flammable</i> F)</p>	 <p>GHS03 (<i>oxidising</i> O)</p>	 <p>GHS05 (<i>corrosive</i> C)</p>
 <p>GHS06 (<i>acutely toxic</i> T)</p>	 <p>GHS07 (<i>moderate hazard</i> MH)</p>	 <p>GHS08 (<i>health hazard</i> HH)</p>
	 <p>GHS09 (<i>hazardous to the aquatic environment</i> N)</p>	

Worksheet

Aim

To investigate the magnitude of cell potentials using a variety of half-cells, and to investigate the effect of change of concentration on the cell potential.

Safety:

- Wear eye protection.
- 1 mol dm⁻³ copper(II) sulfate **[H] [N]**
- 1 mol dm⁻³ zinc sulfate **[H][N]**
- 1 mol dm⁻³ iron(II) sulfate **[H]**
- 1 mol dm⁻³ iron(III) chloride **[H]**
- saturated aqueous potassium nitrate(V) **[O]**

Hazard symbols

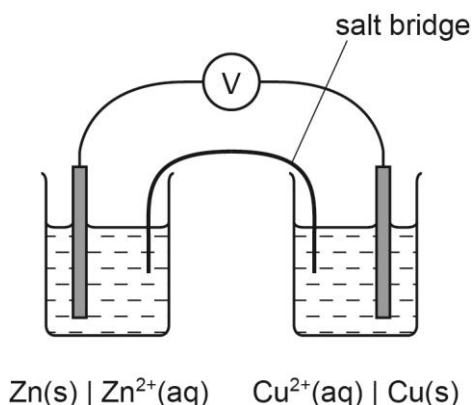
H = harmful or irritating substance

O = oxidising substance

N = harmful to the environment

Method A

1. Pour a small volume of saturated aqueous potassium nitrate(V) into the 250 cm³ beaker. Cut strips of filter paper and soak them in the potassium nitrate. These will be your salt bridges.
2. Use a crocodile clip to grip a strip of clean copper metal in a 100 cm³ (or smaller) beaker. Add 1 mol dm⁻³ copper(II) sulfate to the beaker until the solution covers some of the copper metal. This is your Cu(s) | Cu²⁺(aq) half-cell.
3. Use a crocodile clip to grip a strip of clean zinc metal in a small beaker. Add 1 mol dm⁻³ zinc sulfate to the beaker until the solution covers some of the zinc metal. This is your Zn(s) | Zn²⁺(aq) half-cell.
4. Place the two beakers from 2 and 3 close together. Use connecting wires to connect the metal strips to either side of a high resistance d.c. voltmeter or a multimeter adjusted to the appropriate range.
5. Use forceps to remove a strip of filter paper from the saturated aqueous potassium nitrate and position it so one end is in the aqueous copper(II) sulfate and the other end is in the aqueous zinc sulfate.



Worksheet, *continued*

6. Take a voltmeter reading as soon as the circuit is complete. Record your reading. Record which half-cell is attached to the positive terminal of the voltmeter.
7. Disconnect the voltmeter and discard the salt bridge. Do not dismantle your half-cells.
8. Use a crocodile clip to grip a silver electrode (a silver wire or a volunteer's ring are suitable) in a small beaker. Add 0.05 mol dm^{-3} silver nitrate to the beaker until the solution covers some of the silver metal. This is your $\text{Ag(s)} | \text{Ag}^+(\text{aq})$ half-cell. Place it close to the $\text{Cu(s)} | \text{Cu}^{2+}(\text{aq})$ half-cell.
(If a volunteer's ring is used, make sure it is washed well with water before it is worn again.)
9. Use the connecting wires to connect the metal strips to either side of the high resistance d.c. voltmeter or multimeter.
10. Use forceps to remove a strip of filter paper from the saturated aqueous potassium nitrate(V) and position it so one end is in the aqueous copper(II) sulfate and the other end is in the aqueous silver nitrate.
11. Take a voltmeter reading as soon as the circuit is complete. Record your reading. Record which half-cell is attached to the positive terminal of the voltmeter.
12. Disconnect the voltmeter and discard the salt bridge.
13. Use your voltmeter readings to **predict** the cell potential if the half-cells used are $\text{Ag(s)} | \text{Ag}^+(\text{aq})$ and $\text{Zn(s)} | \text{Zn}^{2+}(\text{aq})$. Which half-cell should be connected to the positive terminal?
14. Set up the cell using a fresh salt bridge. Record the new voltmeter reading and compare it with your predicted value.
15. Use a 25 cm^3 measuring cylinder to place equal volumes of aqueous iron(II) sulfate and aqueous iron(III) chloride in a small beaker. Dip a platinum electrode into the solution. This is the $\text{Pt} | \text{Fe}^{2+}(\text{aq}), \text{Fe}^{3+}(\text{aq})$ half-cell.
16. Connect the $\text{Pt} | \text{Fe}^{2+}(\text{aq}), \text{Fe}^{3+}(\text{aq})$ half cell with one of the others. Record the new voltmeter reading and which half-cell is connected to the positive terminal.
17. Use other available solutions with the appropriate metal strips or electrodes to investigate which half-cell is the best provider of electrons to the external circuit. (The substances available may include magnesium ribbon, 1 mol dm^{-3} magnesium sulfate, 0.02 mol dm^{-3} iodine in approx. 0.05 mol dm^{-3} potassium iodide.)

Results

Record **all** your observations.

Tabulate your results clearly, showing which is the positive half of the cell and recording the value of the cell pd / V.

All voltmeter readings should be shown to the same number of decimal places.

Worksheet, *continued*

Interpretation and evaluation

1. Comment on the accuracy of your predicted value compared with the experimental value you obtained for the $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Ag}^{+}(\text{aq}) \mid \text{Ag(s)}$ cell.
2. The data book value for this combination of half-cells is 1.56 V. Suggest why your experimental value is different from the data book value.
3. Which substance of those you tested has the greatest ability to donate electrons to the external circuit?
4. Which substance of those you tested would you expect to be the best oxidising agent?

Method B

The Nernst equation is given as:

$$E = E^{\ominus} + (0.059/z) \log \frac{[\text{reduced species}]}{[\text{oxidised species}]}$$

where E is the cell emf, E^{\ominus} is the standard cell emf and z is the difference in charge between the reduced and oxidised species.

You should write a plan for experiments you could carry out that will test the Nernst equation.

- You may use any of the apparatus, half-cells and results from method A. You may also use distilled water, solid potassium iodide and other normal laboratory apparatus you may require.
- Make sure that your plan includes the concentrations of the solutions you will use and how you will prepare solutions of these concentrations.
- You should make it clear in your plan what you are keeping the same (controlled variables) and what you are changing (independent variable).

Carry out your planned experiments.

Results

Record and tabulate the results of your experiments.

Tabulate your results clearly, showing which is the positive half of the cell and recording the value of the cell pd / V.

All voltmeter readings should be shown to the same number of decimal places.

Interpretation and evaluation

1. Calculate the theoretical E_{cell} values for your experiments.
2. Compare your experimental values for E_{cell} with those you calculated. Do your results support the Nernst equation?

Cambridge Assessment International Education
The Triangle Building, Shaftesbury Road, Cambridge, CB2 8EA
t: +44 1223 553554 f: +44 1223 553558
e: info@cambridgeinternational.org www.cambridgeinternational.org

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