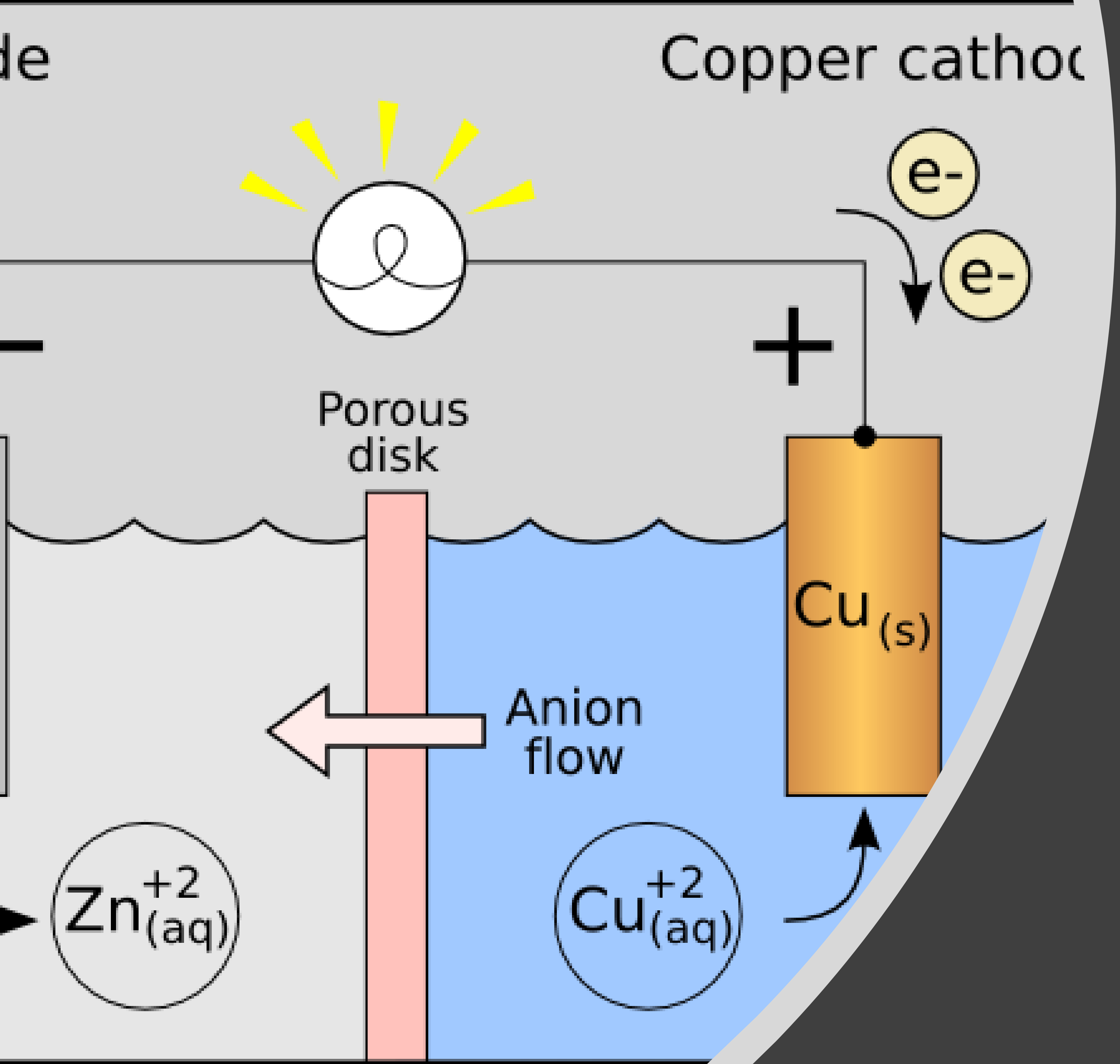


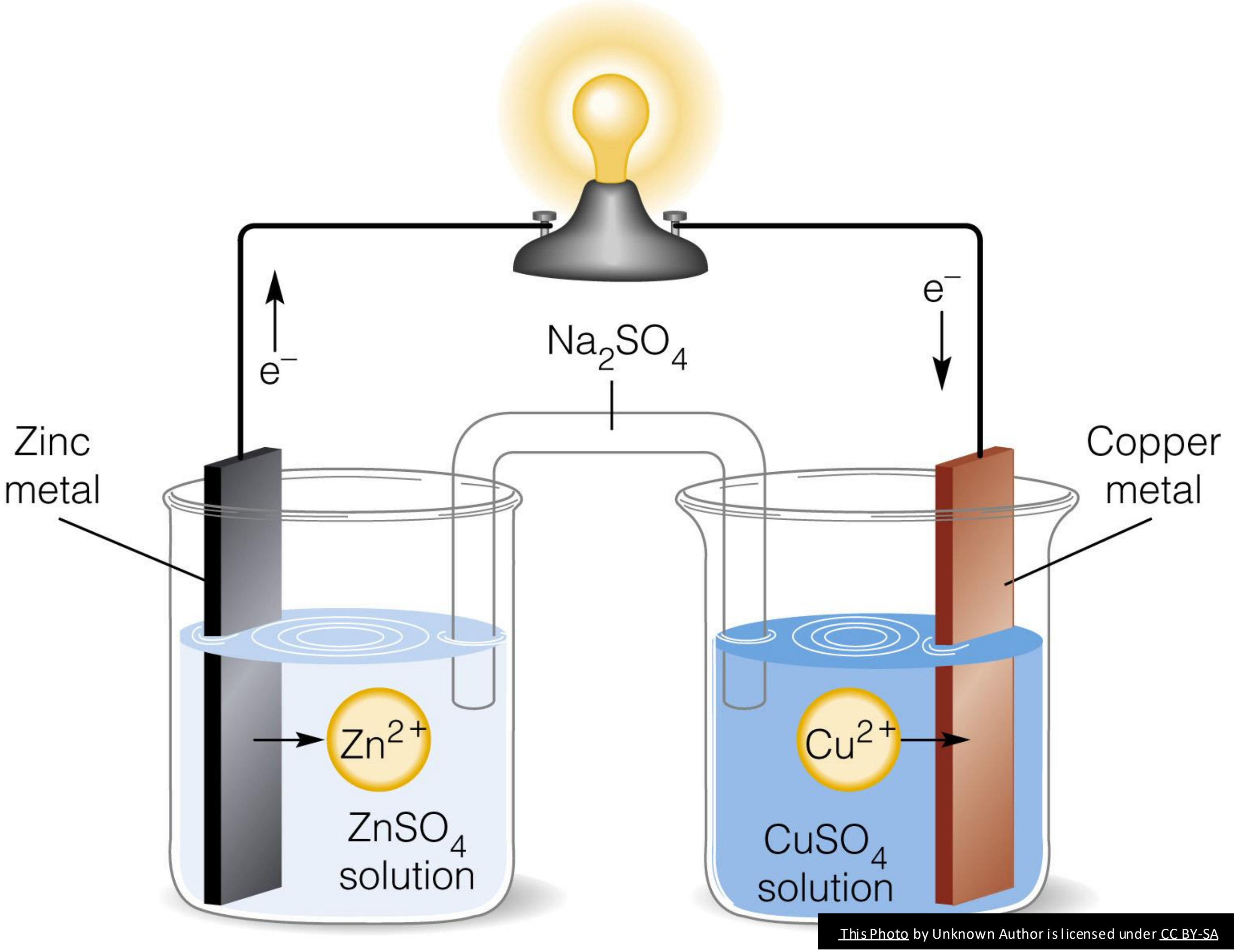
Galvanic cells

Electrochemical reactions

Grade 12

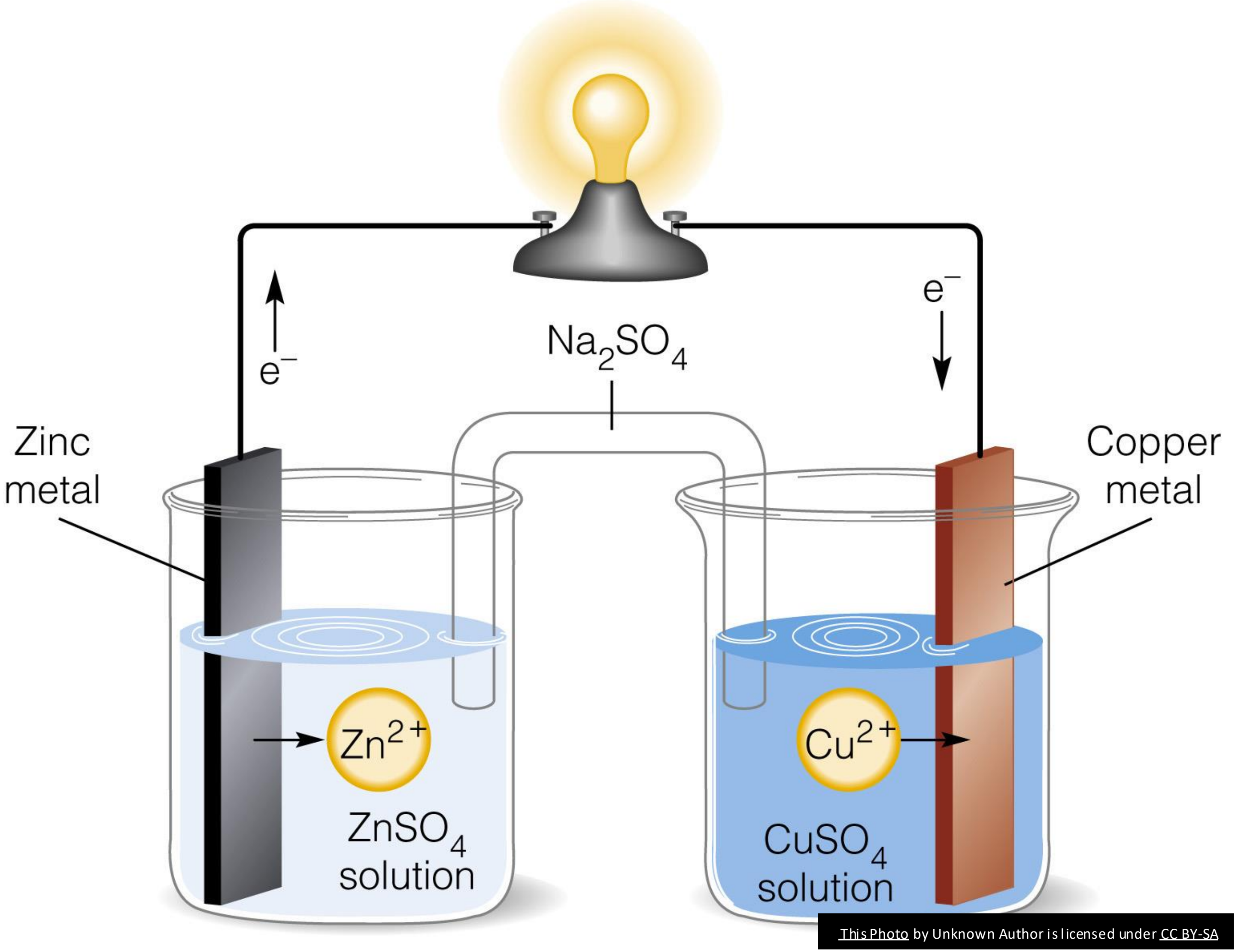
Galvanic cells

- Spontaneous chemical reaction
- Chemical energy converted to electrical energy
- Electrical current generated
- E^{θ}_{cell} is positive



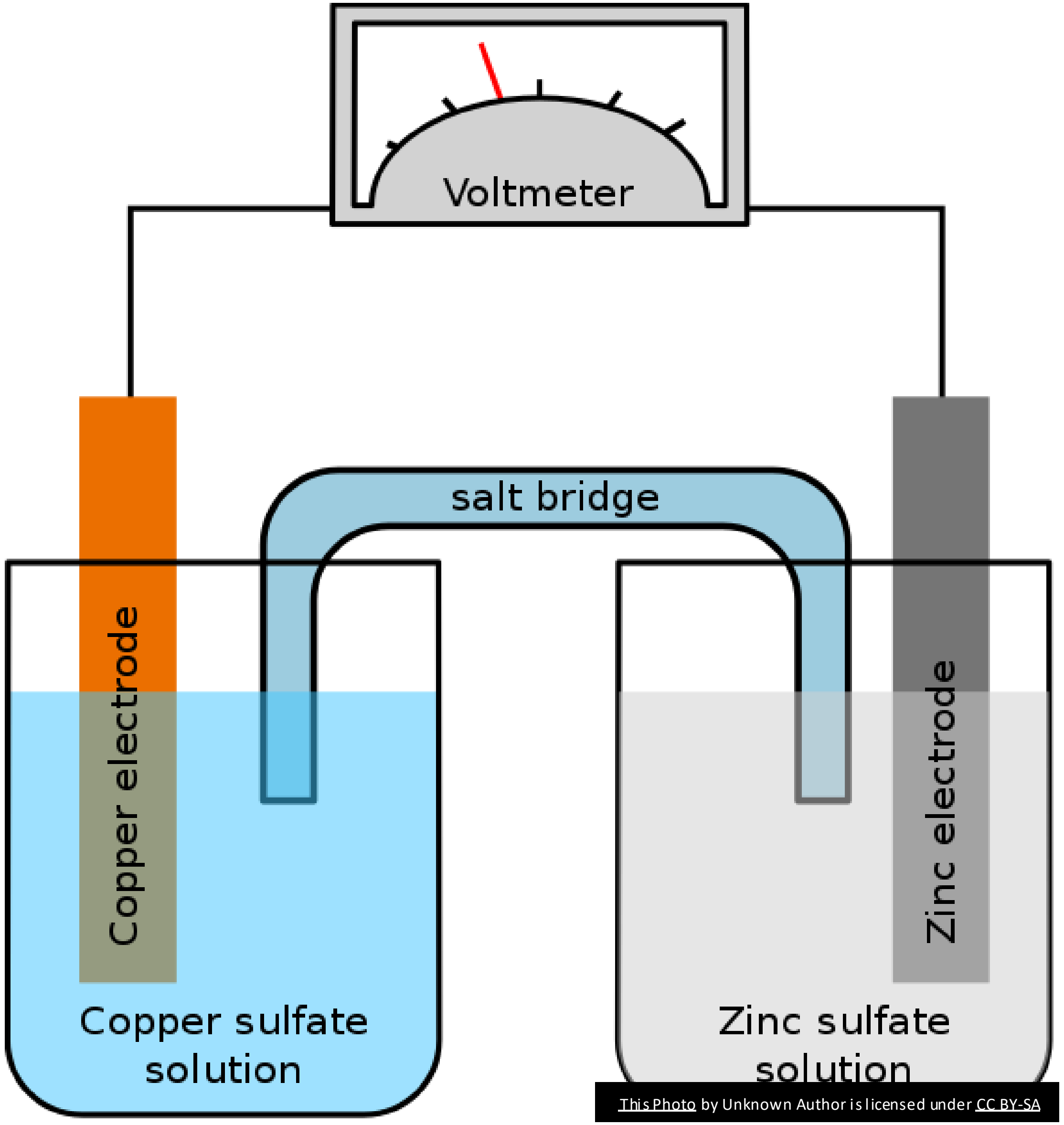
Galvanic cells

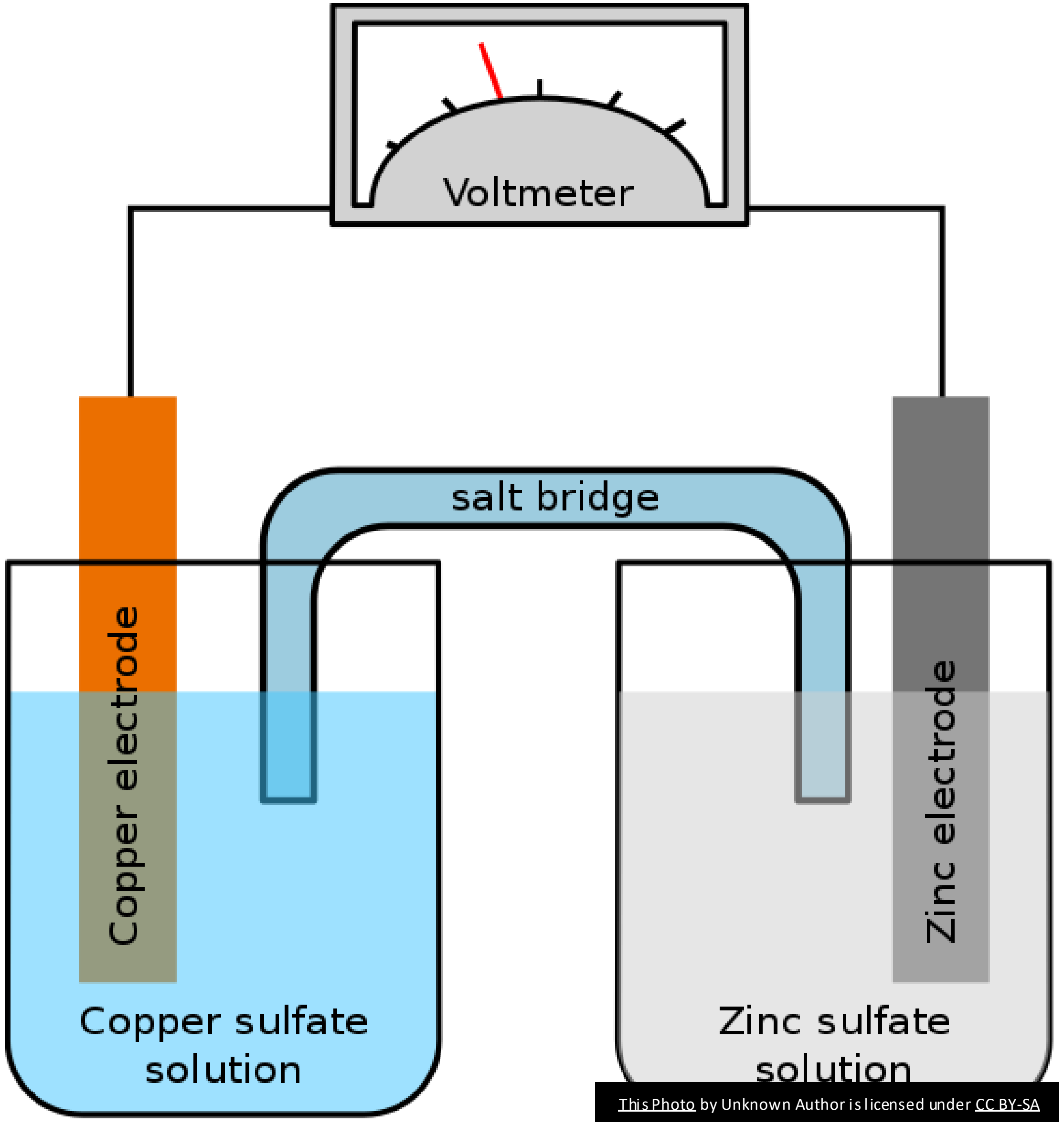
- Anode:
    - Oxidation
    - Negative electrode
    - Decrease in mass
- Cathode
    - Reduction
    - Positive electrode
    - Increase in mass



Functions of a salt bridge

- Separates the electrolytes
- Completes the circuit
- Provides a path through which ions can move to ensure neutrality of the electrolytes
- Maintains electrical neutrality of the electrolyte solutions

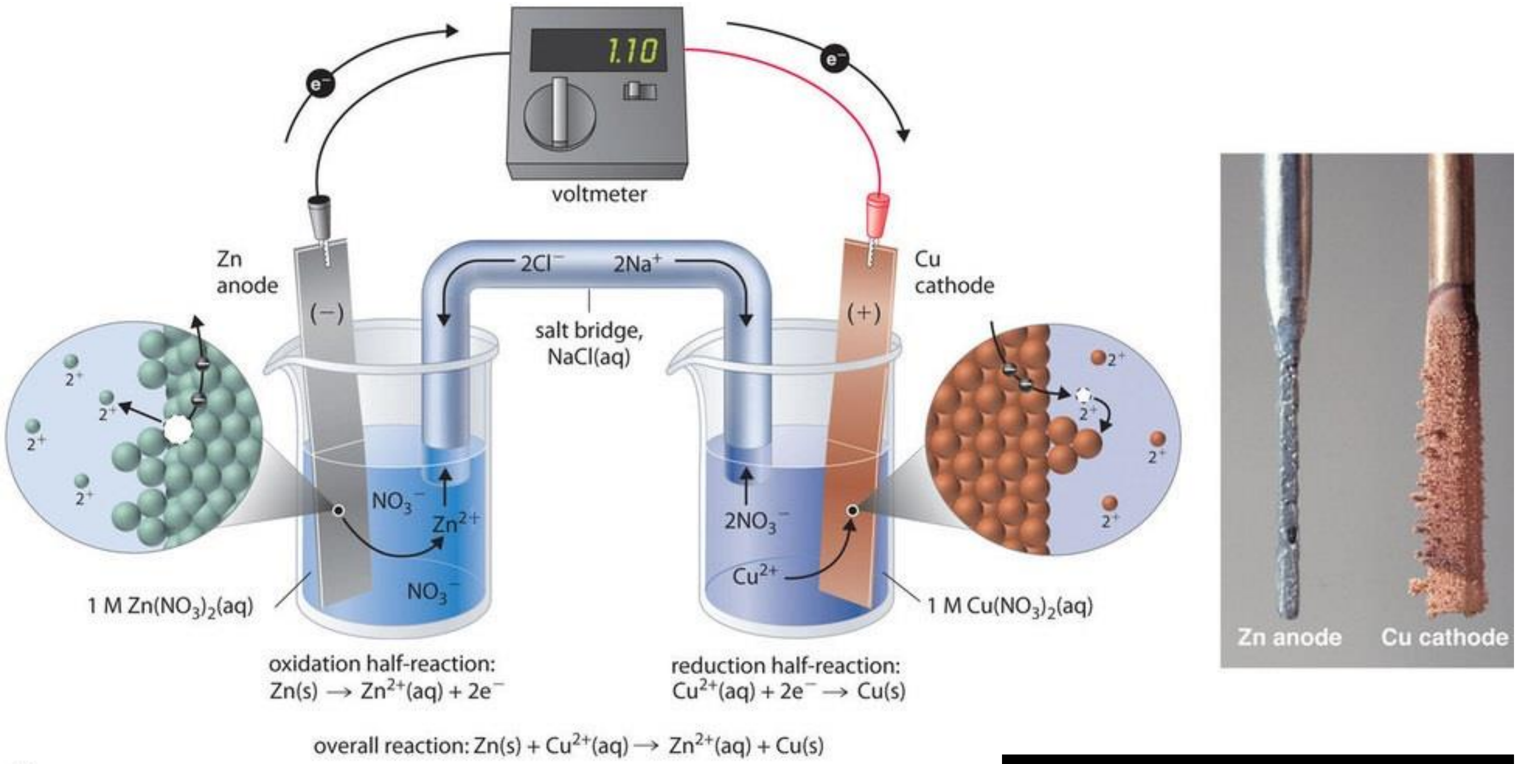




Functions of a salt bridge

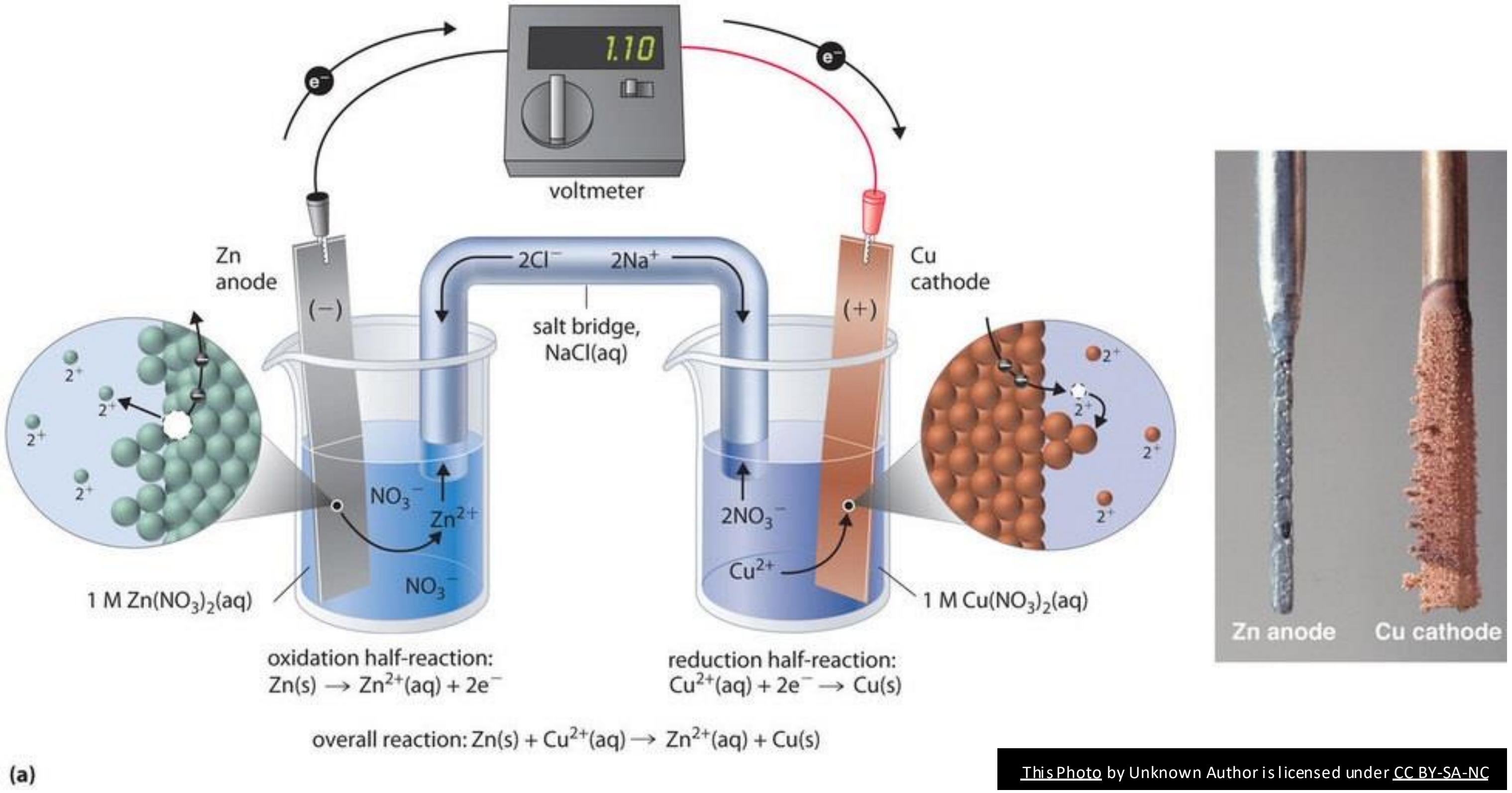
- Usually contains saturated ionic solution
- Concentrated solution reduces internal resistance
- Examples of ionic solutions:
  - NaCl
  - KCl
  - KNO_3
  - Na_2SO_4

Zinc-copper cell



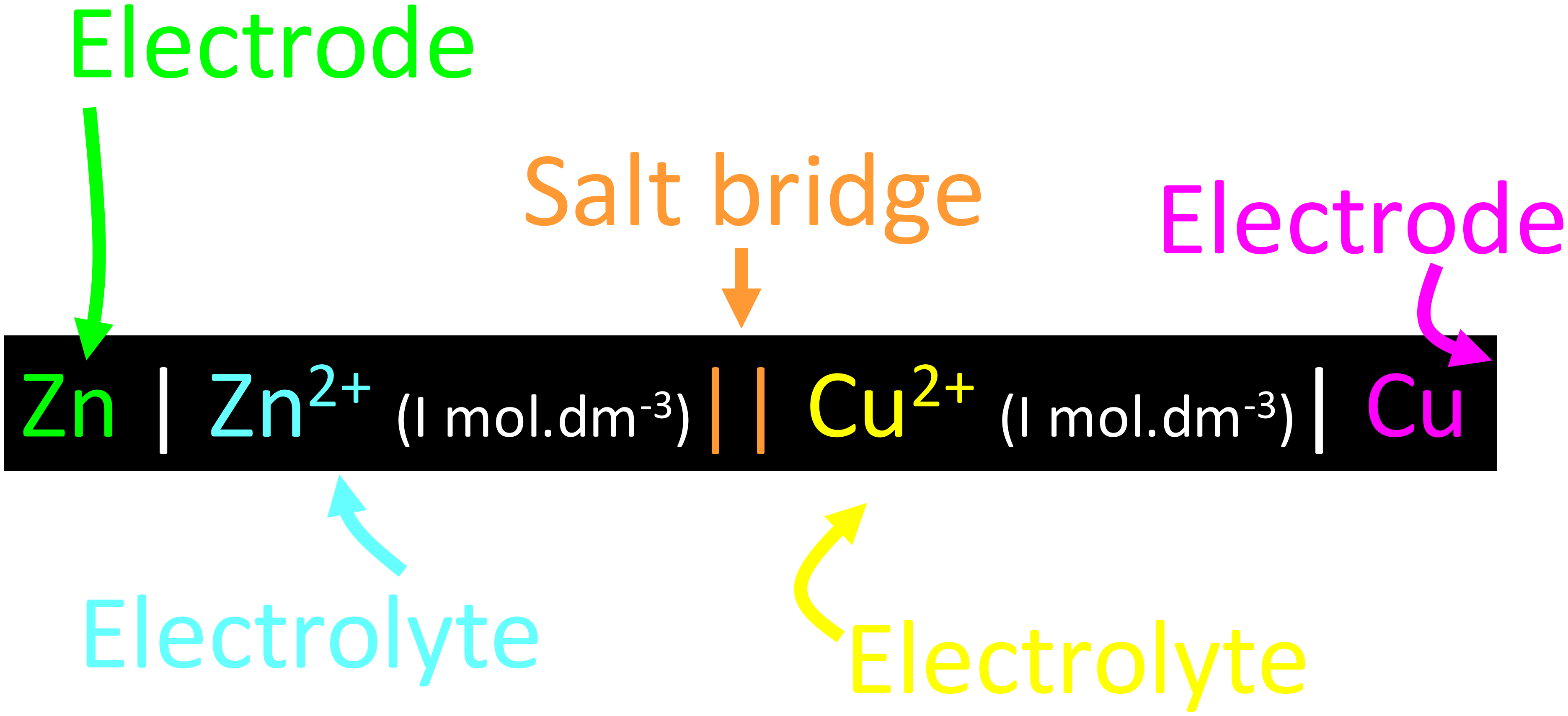
(a)

Zinc-copper cell: cell notation



$Zn \mid Zn^{2+}$ (l mol.dm^{-3}) $\|$ Cu^{2+} (l mol.dm^{-3}) $\mid Cu$

Cell notation



Electrode

Salt bridge

Electrode

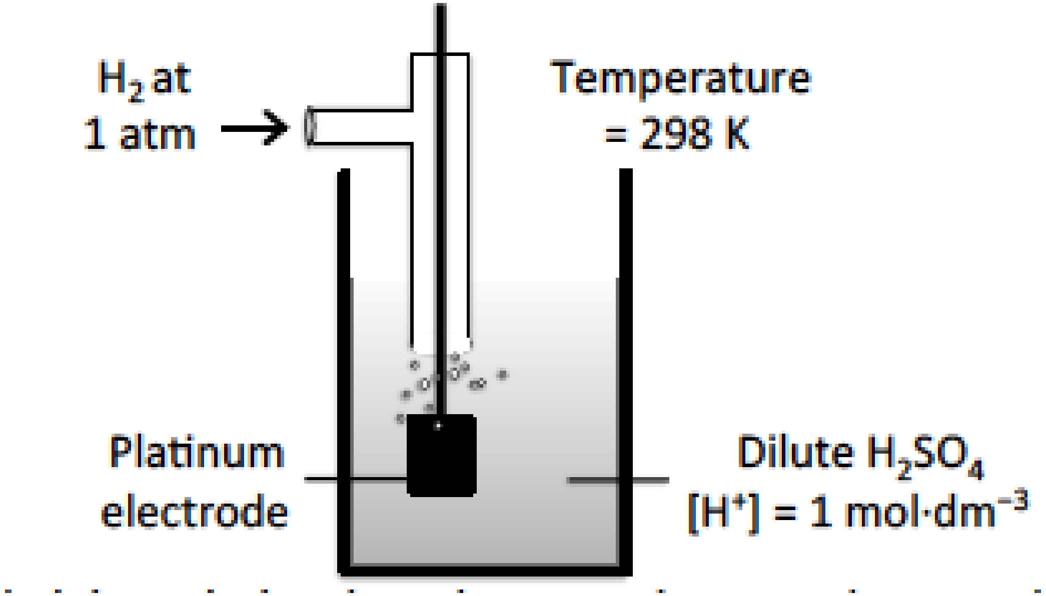
$Zn \mid Zn^{2+}$ (I mol.dm^{-3}) $\mid\mid$ Cu^{2+} (I mol.dm^{-3}) $\mid$ Cu

Electrolyte

Electrolyte

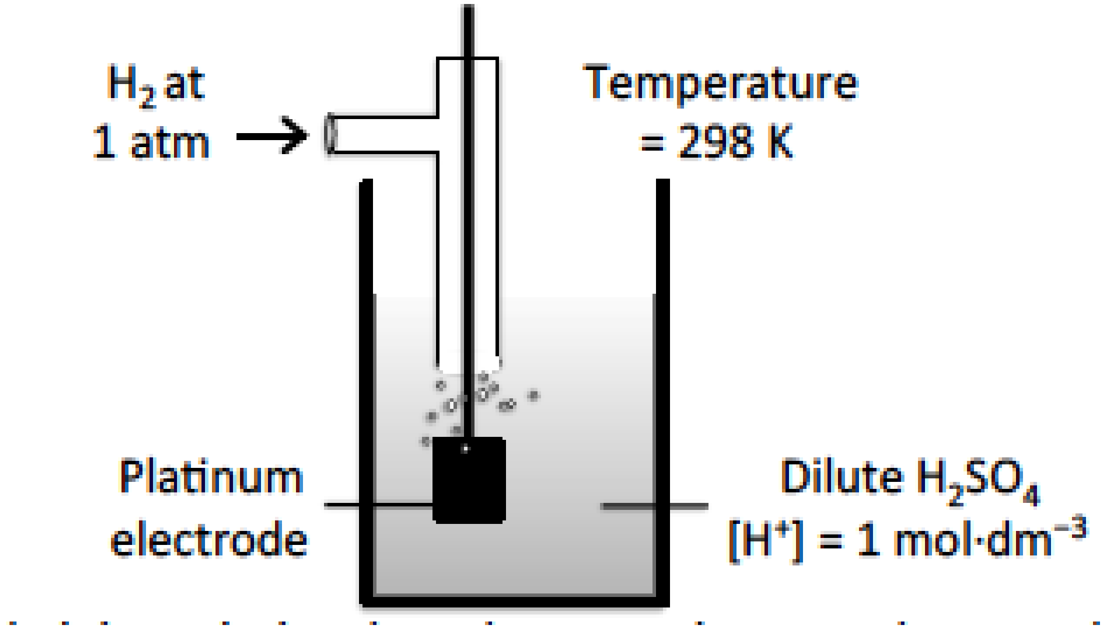
Standard hydrogen electrode

- Hydrogen electrode (H_2) has been chosen as the reference electrode (half cell)
- We compare other half cells to it
- It has a standard electrode potential of 0,00 V
- H_2 is bubbled through a dilute acid electrolyte over a platinum electrode
- Platinum is inert/ inactive
- The Pt absorbs H_2 gas onto its surface, enabling the H_2 to be in contact with the ions in the solution



Standard hydrogen electrode

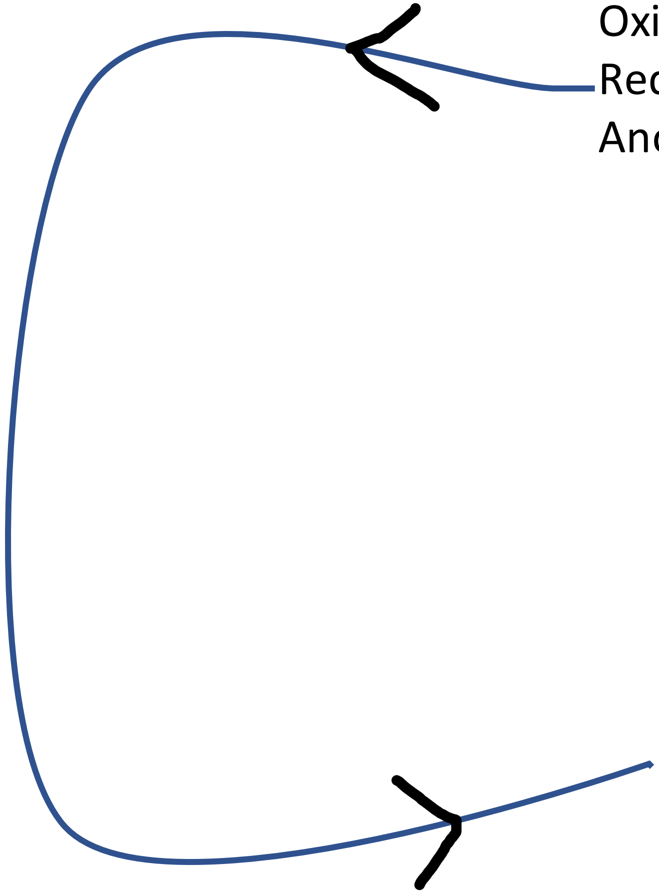
- Standard conditions:
    - Temperature of 25°C or 298K
    - Concentration of solutions of 1 $mol.dm^{-3}$
    - Pressure of 1 atm (only applicable to gasses)



- Cell notation for hydrogen half-cell:

$$Pt \mid H_2\,(g) \mid H^+\,(aq)\ (I\ mol.dm^{-3})$$

Table 4B



Oxidation half reaction
Reducing agent
Anode (-)

Reduction half reaction
Oxidizing agent
Cathode (+)

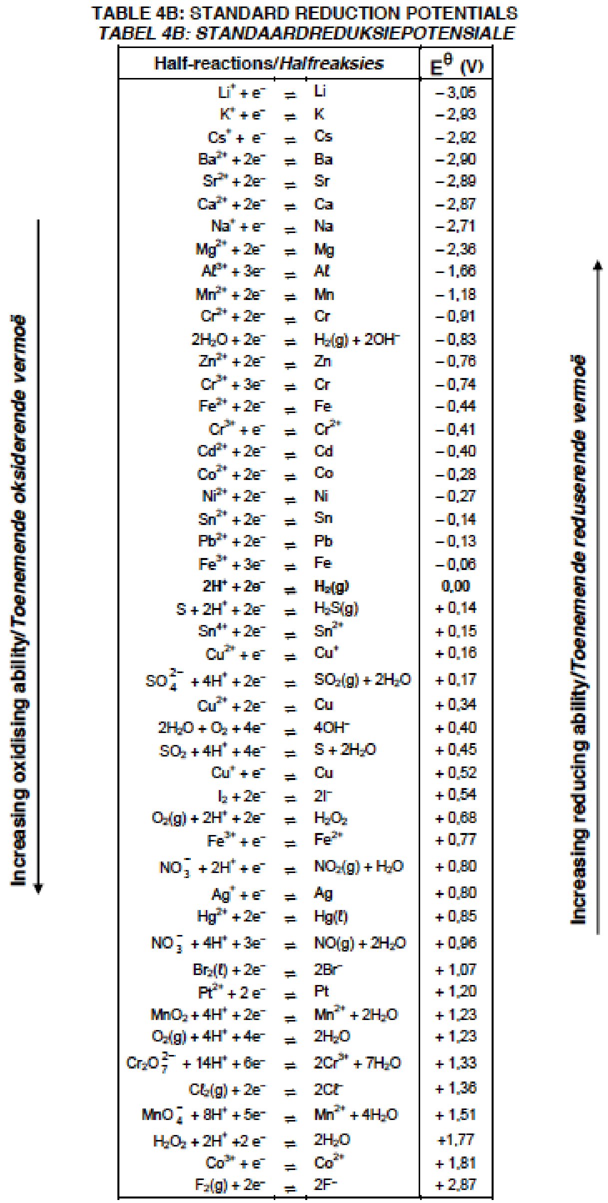


TABLE 4B: STANDARD REDUCTION POTENTIALS
TABEL 4B: STANDAARDREDUKSIEPOTENSIALE

Half-reactions/*Halfreaksies*			E^θ (V)
$Li^+ + e^-$	$\rightleftharpoons$	Li	− 3,05
$K^+ + e^-$	$\rightleftharpoons$	K	− 2,93
$Cs^+ + e^-$	$\rightleftharpoons$	Cs	− 2,92
$Ba^{2+} + 2e^-$	$\rightleftharpoons$	Ba	− 2,90
$Sr^{2+} + 2e^-$	$\rightleftharpoons$	Sr	− 2,89
$Ca^{2+} + 2e^-$	$\rightleftharpoons$	Ca	− 2,87
$Na^+ + e^-$	$\rightleftharpoons$	Na	− 2,71
$Mg^{2+} + 2e^-$	$\rightleftharpoons$	Mg	− 2,36
$A\ell^{3+} + 3e^-$	$\rightleftharpoons$	$A\ell$	− 1,66
$Mn^{2+} + 2e^-$	$\rightleftharpoons$	Mn	− 1,18
$Cr^{2+} + 2e^-$	$\rightleftharpoons$	Cr	− 0,91
$2H_2O + 2e^-$	$\rightleftharpoons$	$H_2(g) + 2OH^-$	− 0,83
$Zn^{2+} + 2e^-$	$\rightleftharpoons$	Zn	− 0,76
$Cr^{3+} + 3e^-$	$\rightleftharpoons$	Cr	− 0,74
$Fe^{2+} + 2e^-$	$\rightleftharpoons$	Fe	− 0,44
$Cr^{3+} + e^-$	$\rightleftharpoons$	Cr^{2+}	− 0,41
$Cd^{2+} + 2e^-$	$\rightleftharpoons$	Cd	− 0,40
$Co^{2+} + 2e^-$	$\rightleftharpoons$	Co	− 0,28
$Ni^{2+} + 2e^-$	$\rightleftharpoons$	Ni	− 0,27
$Sn^{2+} + 2e^-$	$\rightleftharpoons$	Sn	− 0,14
$Pb^{2+} + 2e^-$	$\rightleftharpoons$	Pb	− 0,13
$Fe^{3+} + 3e^-$	$\rightleftharpoons$	Fe	− 0,06
$\mathbf{2H^+ + 2e^-}$	$\rightleftharpoons$	$\mathbf{H_2(g)}$	**0,00**
$S + 2H^+ + 2e^-$	$\rightleftharpoons$	$H_2S(g)$	+ 0,14
$Sn^{4+} + 2e^-$	$\rightleftharpoons$	Sn^{2+}	+ 0,15
$Cu^{2+} + e^-$	$\rightleftharpoons$	Cu^+	+ 0,16
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2(g) + 2H_2O$	+ 0,17
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	Cu	+ 0,34
$2H_2O + O_2 + 4e^-$	$\rightleftharpoons$	$4OH^-$	+ 0,40
$SO_2 + 4H^+ + 4e^-$	$\rightleftharpoons$	$S + 2H_2O$	+ 0,45
$Cu^+ + e^-$	$\rightleftharpoons$	Cu	+ 0,52
$I_2 + 2e^-$	$\rightleftharpoons$	$2I^-$	+ 0,54
$O_2(g) + 2H^+ + 2e^-$	$\rightleftharpoons$	H_2O_2	+ 0,68
$Fe^{3+} + e^-$	$\rightleftharpoons$	Fe^{2+}	+ 0,77
$NO_3^- + 2H^+ + e^-$	$\rightleftharpoons$	$NO_2(g) + H_2O$	+ 0,80
$Ag^+ + e^-$	$\rightleftharpoons$	Ag	+ 0,80
$Hg^{2+} + 2e^-$	$\rightleftharpoons$	$Hg(\ell)$	+ 0,85
$NO_3^- + 4H^+ + 3e^-$	$\rightleftharpoons$	$NO(g) + 2H_2O$	+ 0,96
$Br_2(\ell) + 2e^-$	$\rightleftharpoons$	$2Br^-$	+ 1,07
$Pt^{2+} + 2e^-$	$\rightleftharpoons$	Pt	+ 1,20
$MnO_2 + 4H^+ + 2e^-$	$\rightleftharpoons$	$Mn^{2+} + 2H_2O$	+ 1,23
$O_2(g) + 4H^+ + 4e^-$	$\rightleftharpoons$	$2H_2O$	+ 1,23
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightleftharpoons$	$2Cr^{3+} + 7H_2O$	+ 1,33
$C\ell_2(g) + 2e^-$	$\rightleftharpoons$	$2C\ell^-$	+ 1,36
$MnO_4^- + 8H^+ + 5e^-$	$\rightleftharpoons$	$Mn^{2+} + 4H_2O$	+ 1,51
$H_2O_2 + 2H^+ + 2e^-$	$\rightleftharpoons$	$2H_2O$	+1,77
$Co^{3+} + e^-$	$\rightleftharpoons$	Co^{2+}	+ 1,81
$F_2(g) + 2e^-$	$\rightleftharpoons$	$2F^-$	+ 2,87

Increasing oxidising ability/Toenemende oksiderende vermoë

Increasing reducing ability/Toenemende reduserende vermoë

Copper and hydrogen half-cells

The half-cell reactions are as follows:

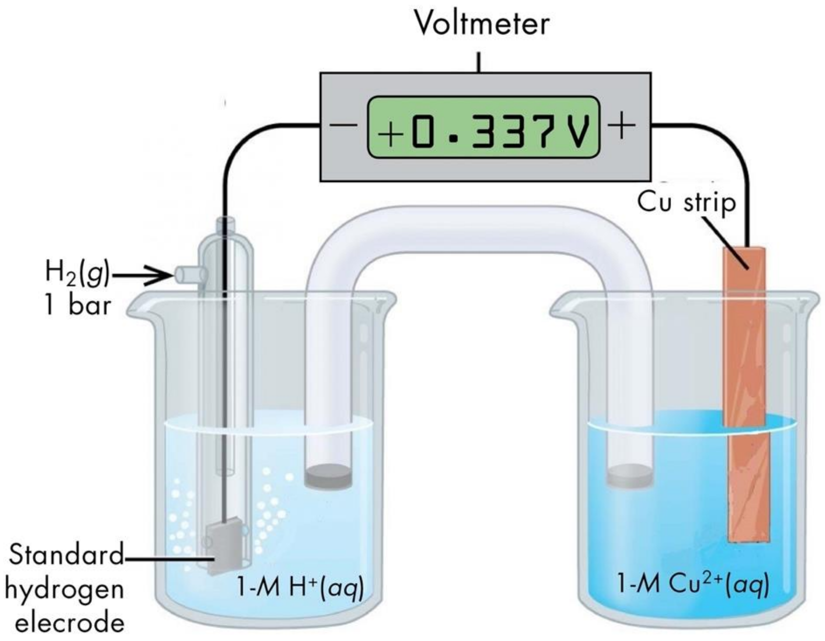
Oxidation half-reaction: $H_2 \rightarrow 2H^+ + 2e^-$

Reduction half-reaction: $Cu^{2+} + 2e^- \rightarrow Cu$

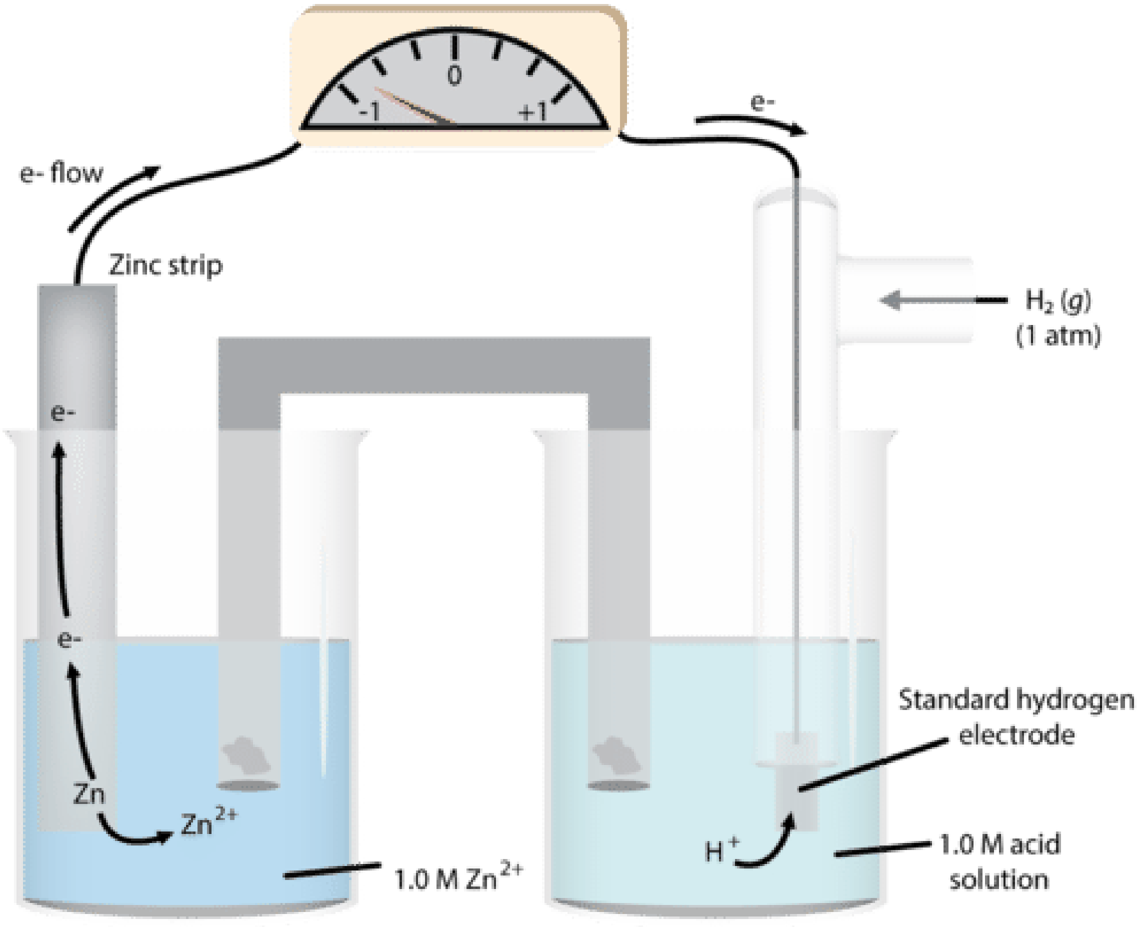
Net cell reaction: $Cu^{2+} + H_2 \rightarrow Cu + 2H^+$

Copper and hydrogen half-cells: Cell notation

$Pt \mid H_2 (g) \mid H^+ (aq) (I\ mol.dm^{-3}) \| Cu^{2+} (aq)(I\ mol.dm^{-3}) \mid Cu(s)$

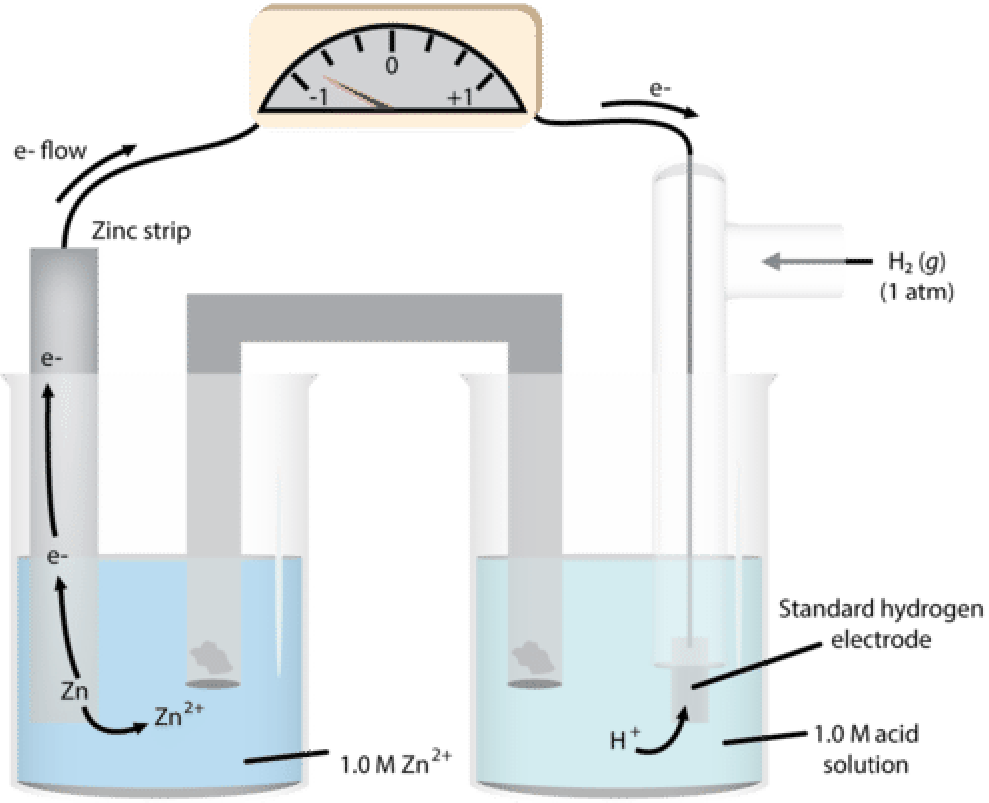


Zinc and hydrogen half-cells: Cell notation



Zinc and hydrogen half-cells: Cell notation

Zn (s) |Zn^{2+} (aq) (I $mol.dm^{-3}$)| | H^+ (aq) (I $mol.dm^{-3}$) | H_2 (g) | Pt

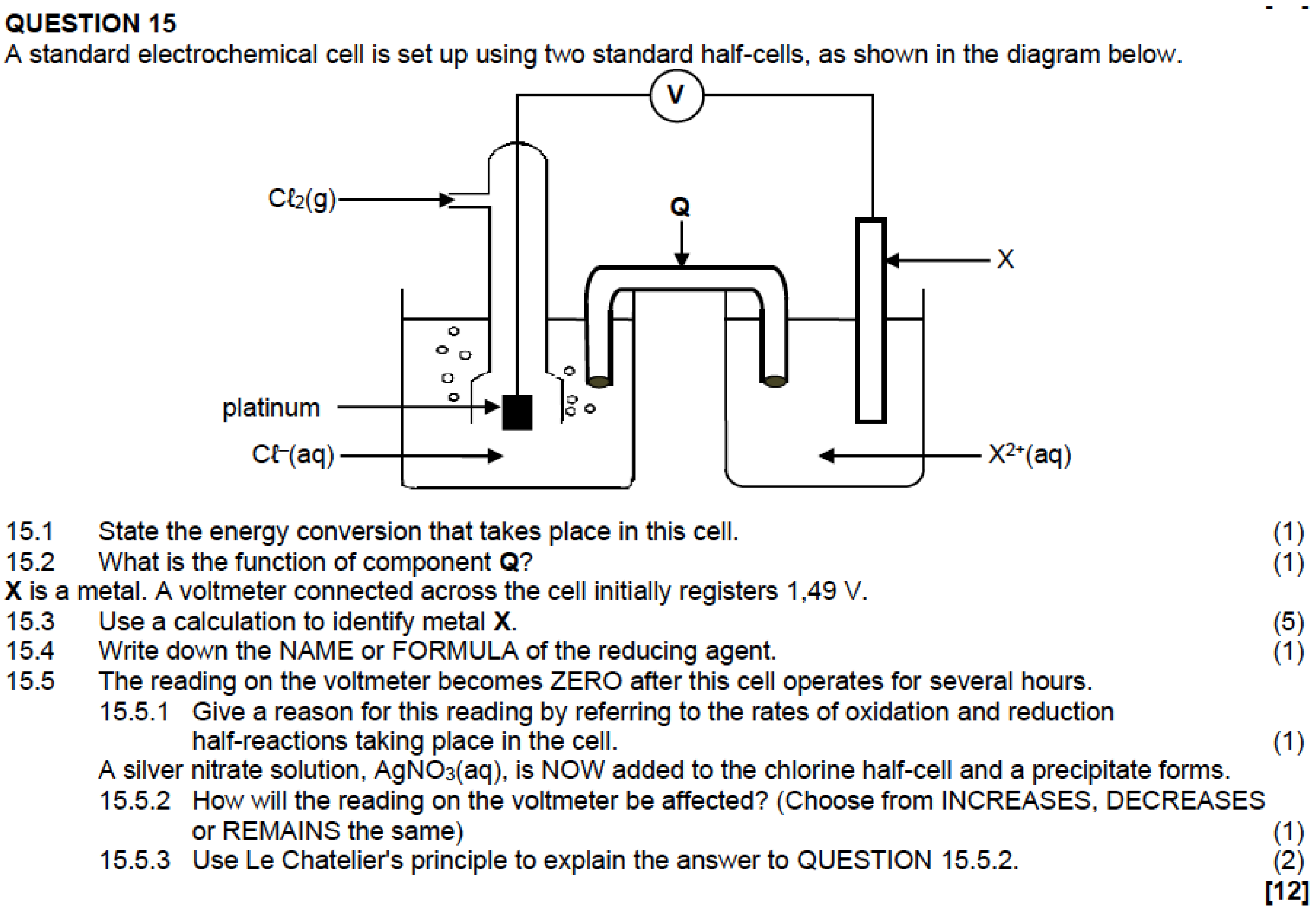


Cell potential

$$E^{\theta}_{cell} = E^{\theta}_{cathode} - E^{\theta}_{anode}$$

- Positive cell potential – spontaneous reaction
- Indicates that the FORWARD reaction of the redox reaction is favoured
- The higher the value, the further the equilibrium moves to the right

Example



QUESTION 15

A standard electrochemical cell is set up using two standard half-cells, as shown in the diagram below.

15.1 State the energy conversion that takes place in this cell. (1)

15.2 What is the function of component **Q**? (1)

X is a metal. A voltmeter connected across the cell initially registers 1,49 V.

15.3 Use a calculation to identify metal **X**. (5)

15.4 Write down the NAME or FORMULA of the reducing agent. (1)

15.5 The reading on the voltmeter becomes ZERO after this cell operates for several hours.

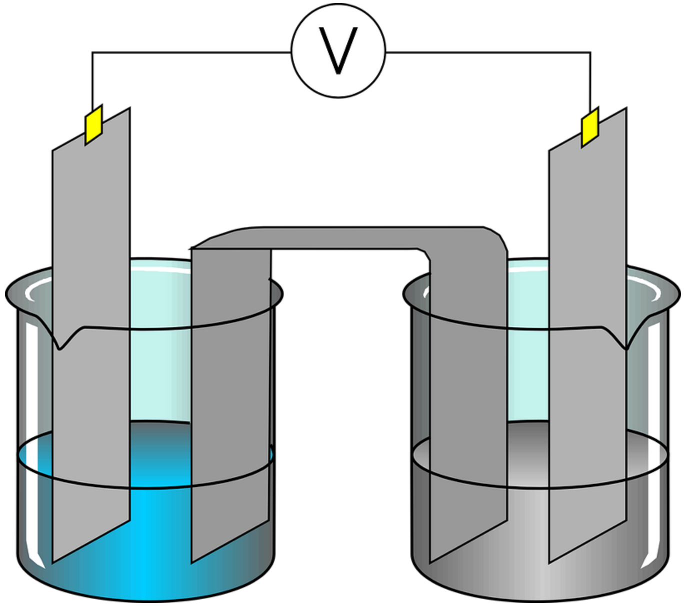
15.5.1 Give a reason for this reading by referring to the rates of oxidation and reduction half-reactions taking place in the cell. (1)

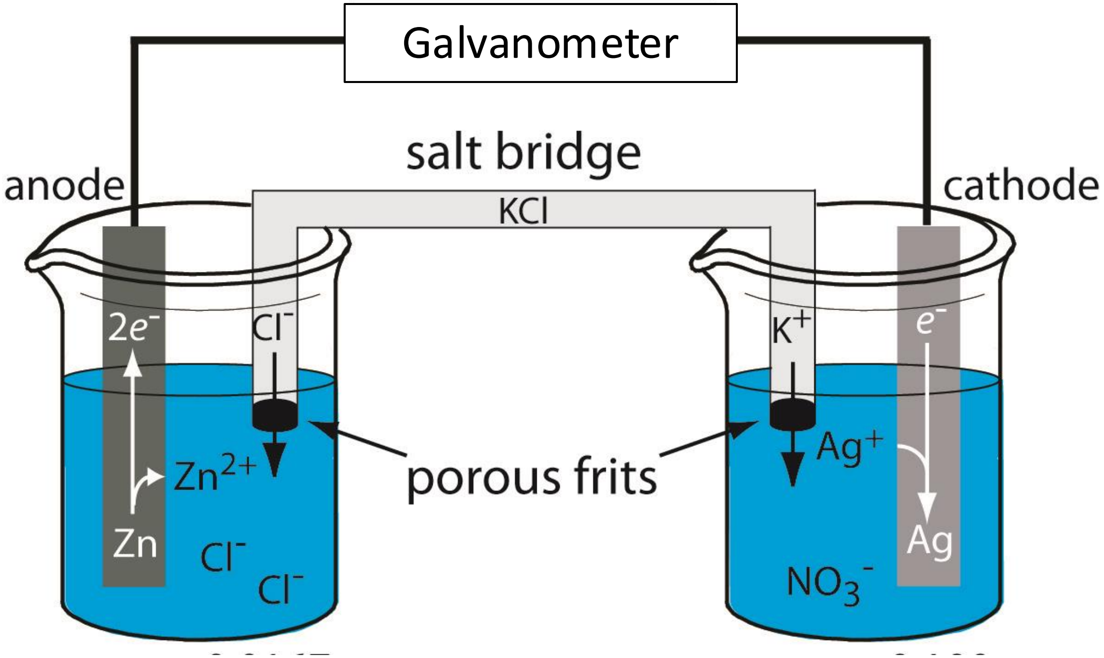
A silver nitrate solution, $AgNO_3(aq)$, is NOW added to the chlorine half-cell and a precipitate forms.

15.5.2 How will the reading on the voltmeter be affected? (Choose from INCREASES, DECREASES or REMAINS the same) (1)

15.5.3 Use Le Chatelier's principle to explain the answer to QUESTION 15.5.2. (2)

[12]

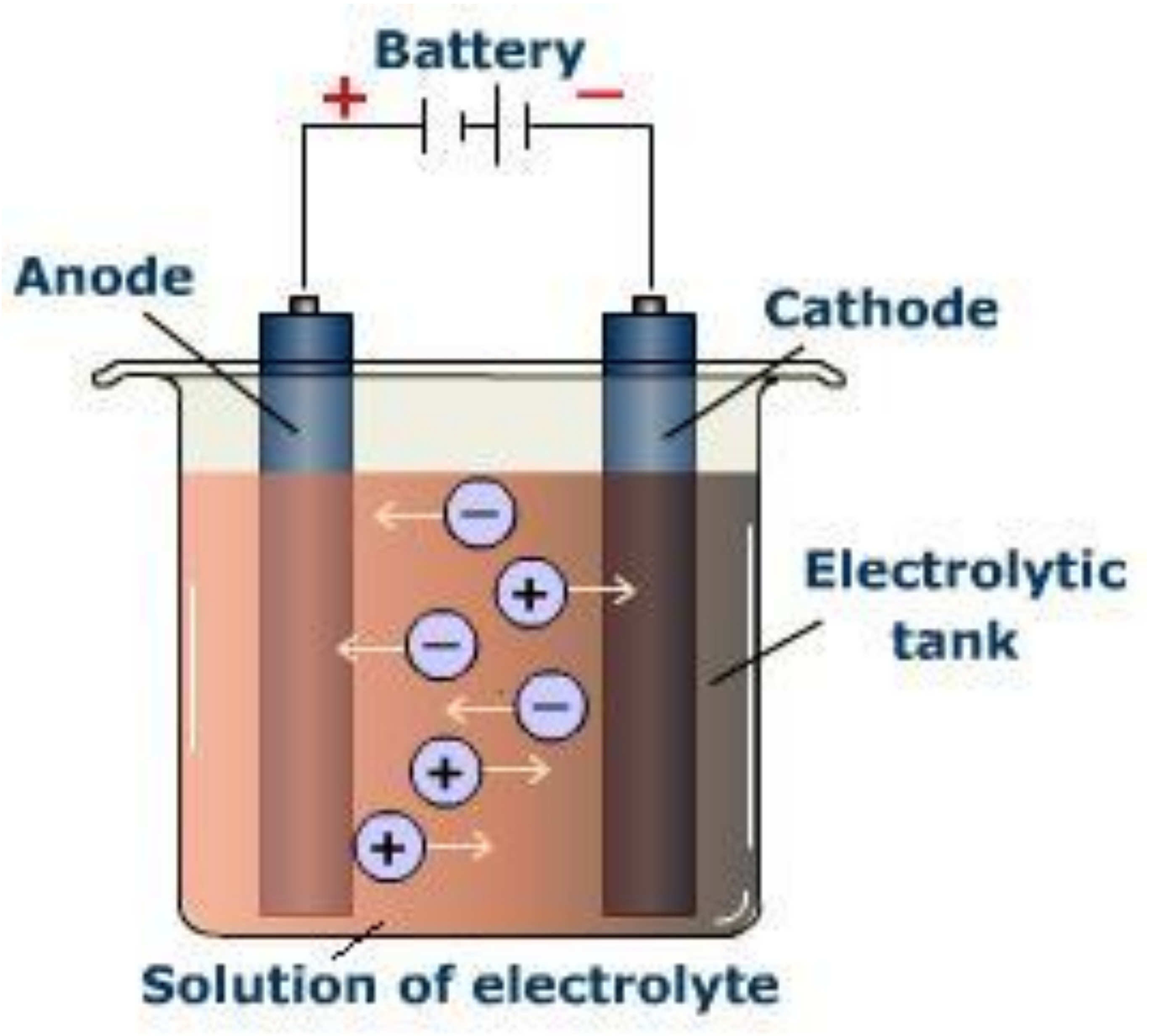

V

Galvanometer
salt bridge
KCl
anode
cathode
2e⁻
Cl⁻
K⁺
e⁻
Zn²⁺
Zn
Cl⁻
Cl⁻
porous frits
Ag⁺
Ag
NO₃⁻

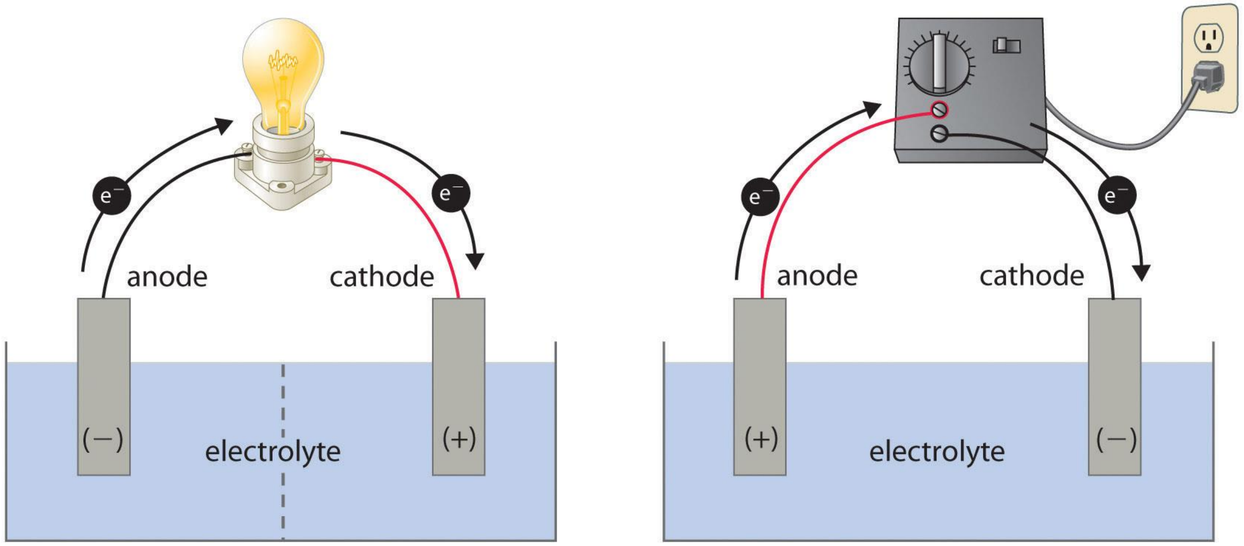
Electrolytic cells

Electrochemical reactions

Grade 12



Galvanic vs Electrolytic cells



GALVANIC CELL

Energy released by spontaneous redox reaction is converted to electrical energy.

Oxidation half-reaction:
$Y \rightarrow Y^+ + e^-$

Reduction half-reaction:
$Z + e^- \rightarrow Z^-$

Overall cell reaction:
$Y + Z \rightarrow Y^+ + Z^-$ $(G < 0)$

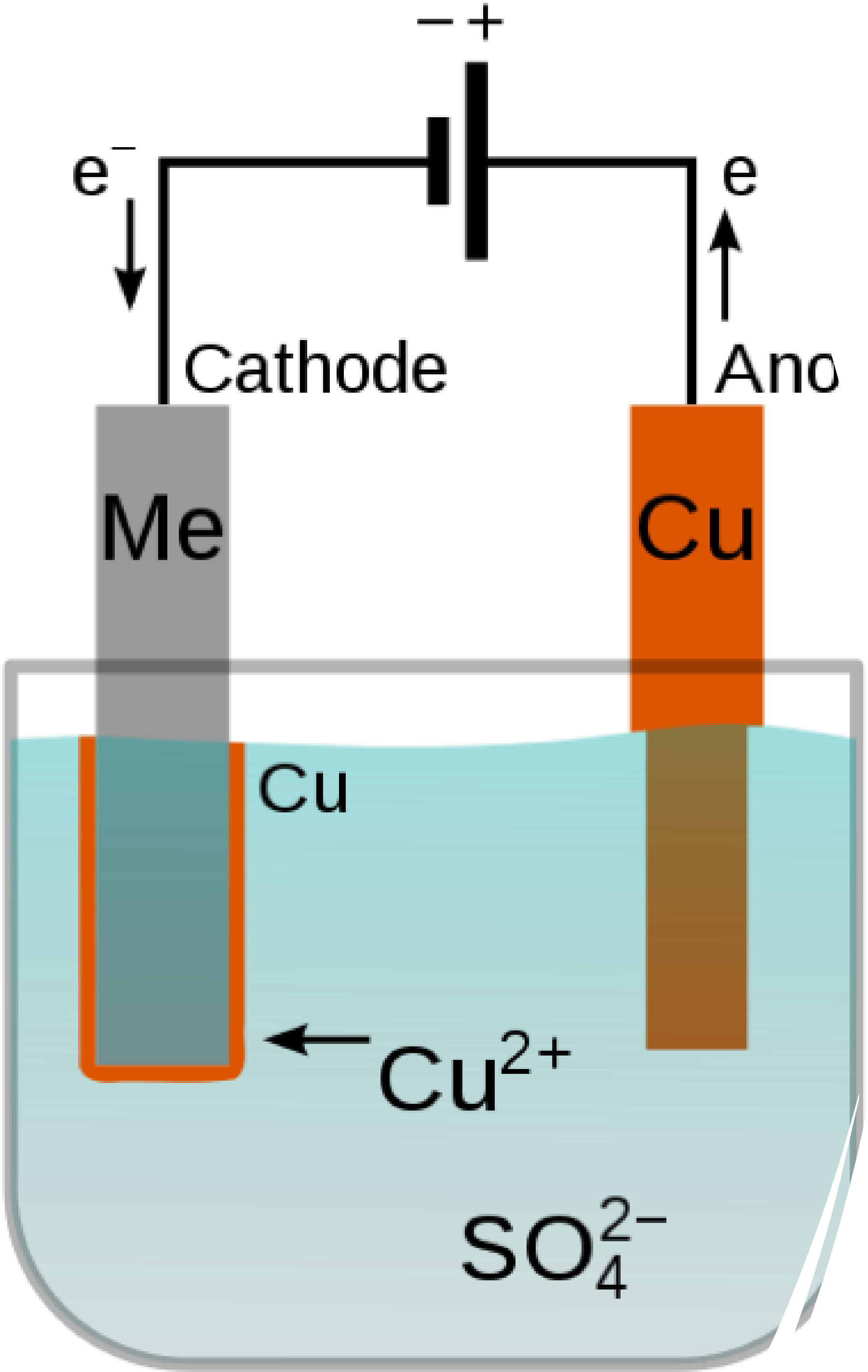
ELECTROLYTIC CELL

Electrical energy is used to drive nonspontaneous redox reaction.

Oxidation half-reaction:
$Z^- \rightarrow Z + e^-$

Reduction half-reaction:
$Y^+ + e^- \rightarrow Y$

Overall cell reaction:
$Y^+ + Z^- \rightarrow Y + Z$ $(G > 0)$



Electrolytic cells

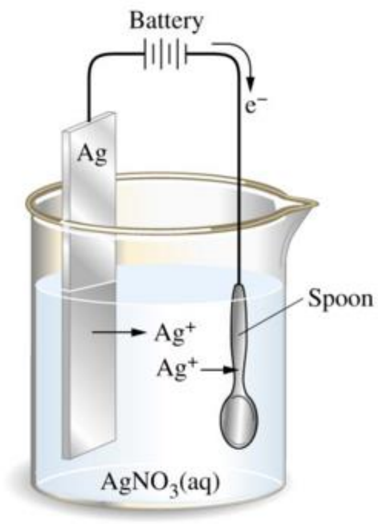
- Non-spontaneous chemical reaction
- Electrical energy converted to chemical energy
- Needs a battery
- Anode is +
- Cathode is -
- E^{θ}_{cell} is negative

Differences between galvanic cells and electrolytic cells

Chemical energy to electrical energy	Electrical energy to chemical energy
Spontaneous	Non-spontaneous
Exothermic	Endothermic
Separate beakers/ electrolytes	Same electrolyte in one beaker
Salt bridge connects two half-cells	No salt bridge
Positive cell potential	Negative cell potential
Anode is negative	**Anode connected to positive** terminal
Cathode is positive	**Cathode connected to negative** terminal

Similarities between galvanic cells and electrolytic cells

- Oxidation at anode
- Negative ions move towards anode
- Reduction at cathode
- Positive ions move towards cathode
- Electrons flow from the anode to cathode
- The mass of the anode decreases
- The mass of the cathode increases



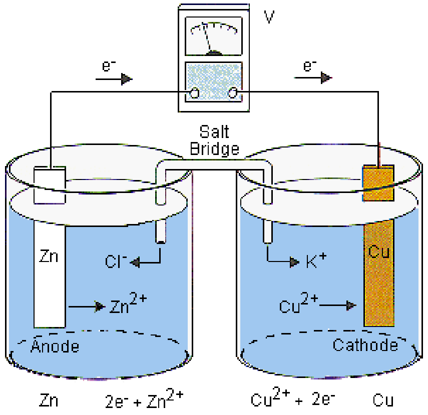
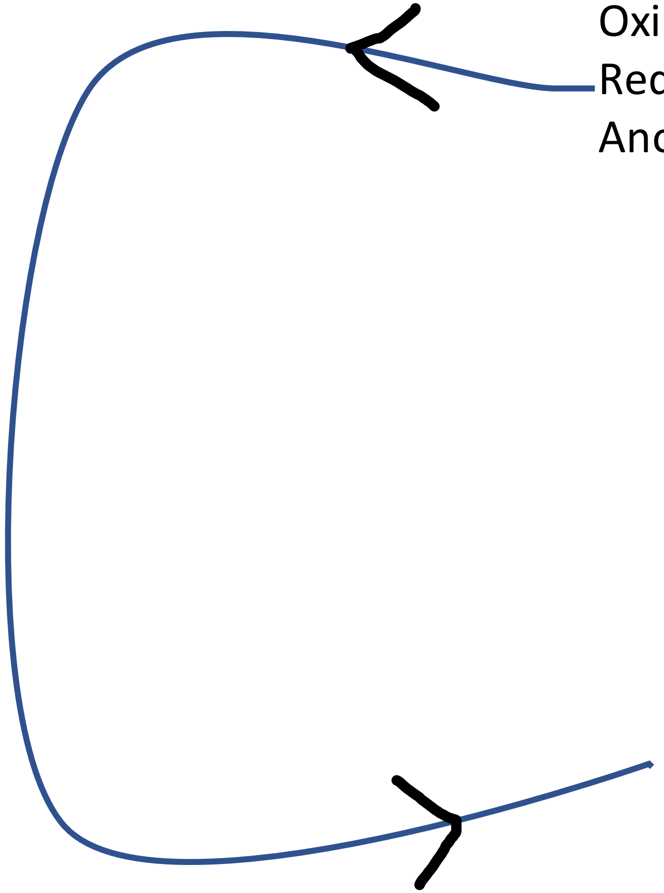


Table 4A



Oxidation half reaction
Reducing agent
Anode (+)

Reduction half reaction
Oxidizing agent
Cathode (-)

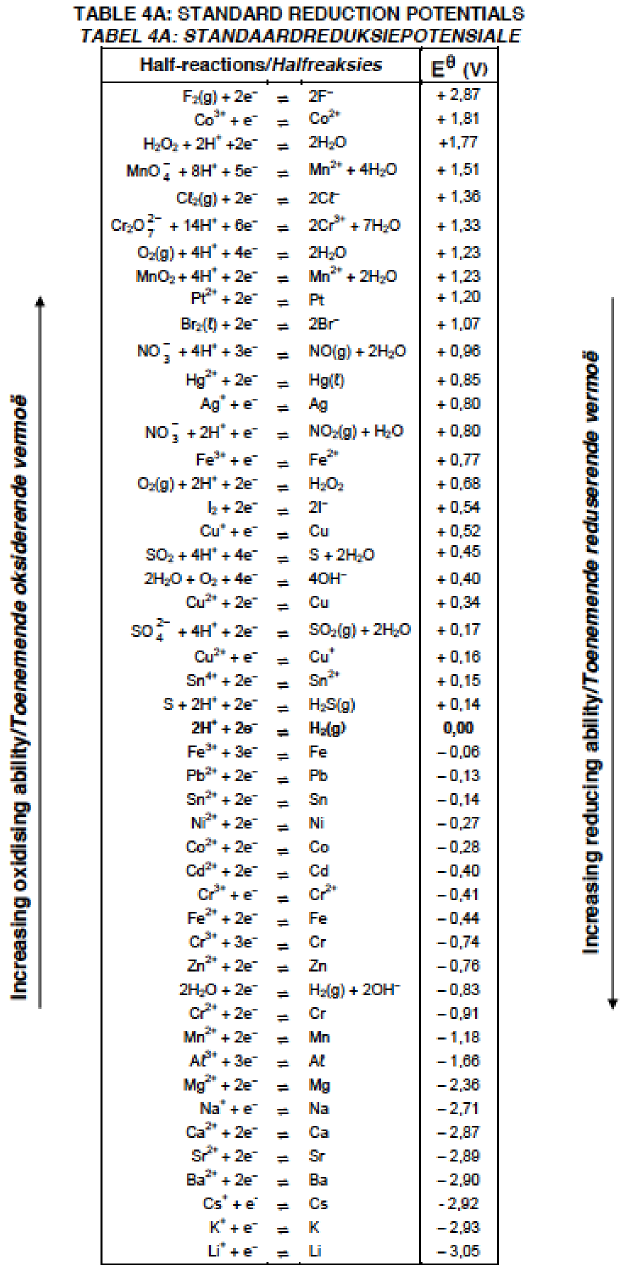


TABLE 4A: STANDARD REDUCTION POTENTIALS
TABEL 4A: STANDAARDREDUKSIEPOTENSIALE

Half-reactions/*Halfreaksies*			**E^θ (V)**
$F_2(g) + 2e^-$	$\rightleftharpoons$	$2F^-$	+ 2,87
$Co^{3+} + e^-$	$\rightleftharpoons$	Co^{2+}	+ 1,81
$H_2O_2 + 2H^+ + 2e^-$	$\rightleftharpoons$	$2H_2O$	+1,77
$MnO_4^- + 8H^+ + 5e^-$	$\rightleftharpoons$	$Mn^{2+} + 4H_2O$	+ 1,51
$C\ell_2(g) + 2e^-$	$\rightleftharpoons$	$2C\ell^-$	+ 1,36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightleftharpoons$	$2Cr^{3+} + 7H_2O$	+ 1,33
$O_2(g) + 4H^+ + 4e^-$	$\rightleftharpoons$	$2H_2O$	+ 1,23
$MnO_2 + 4H^+ + 2e^-$	$\rightleftharpoons$	$Mn^{2+} + 2H_2O$	+ 1,23
$Pt^{2+} + 2e^-$	$\rightleftharpoons$	Pt	+ 1,20
$Br_2(\ell) + 2e^-$	$\rightleftharpoons$	$2Br^-$	+ 1,07
$NO_3^- + 4H^+ + 3e^-$	$\rightleftharpoons$	$NO(g) + 2H_2O$	+ 0,96
$Hg^{2+} + 2e^-$	$\rightleftharpoons$	$Hg(\ell)$	+ 0,85
$Ag^+ + e^-$	$\rightleftharpoons$	Ag	+ 0,80
$NO_3^- + 2H^+ + e^-$	$\rightleftharpoons$	$NO_2(g) + H_2O$	+ 0,80
$Fe^{3+} + e^-$	$\rightleftharpoons$	Fe^{2+}	+ 0,77
$O_2(g) + 2H^+ + 2e^-$	$\rightleftharpoons$	H_2O_2	+ 0,68
$I_2 + 2e^-$	$\rightleftharpoons$	$2I^-$	+ 0,54
$Cu^+ + e^-$	$\rightleftharpoons$	Cu	+ 0,52
$SO_2 + 4H^+ + 4e^-$	$\rightleftharpoons$	$S + 2H_2O$	+ 0,45
$2H_2O + O_2 + 4e^-$	$\rightleftharpoons$	$4OH^-$	+ 0,40
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	Cu	+ 0,34
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2(g) + 2H_2O$	+ 0,17
$Cu^{2+} + e^-$	$\rightleftharpoons$	Cu^+	+ 0,16
$Sn^{4+} + 2e^-$	$\rightleftharpoons$	Sn^{2+}	+ 0,15
$S + 2H^+ + 2e^-$	$\rightleftharpoons$	$H_2S(g)$	+ 0,14
$2H^+ + 2e^-$	$\rightleftharpoons$	**$H_2(g)$**	**0,00**
$Fe^{3+} + 3e^-$	$\rightleftharpoons$	Fe	– 0,06
$Pb^{2+} + 2e^-$	$\rightleftharpoons$	Pb	– 0,13
$Sn^{2+} + 2e^-$	$\rightleftharpoons$	Sn	– 0,14
$Ni^{2+} + 2e^-$	$\rightleftharpoons$	Ni	– 0,27
$Co^{2+} + 2e^-$	$\rightleftharpoons$	Co	– 0,28
$Cd^{2+} + 2e^-$	$\rightleftharpoons$	Cd	– 0,40
$Cr^{3+} + e^-$	$\rightleftharpoons$	Cr^{2+}	– 0,41
$Fe^{2+} + 2e^-$	$\rightleftharpoons$	Fe	– 0,44
$Cr^{3+} + 3e^-$	$\rightleftharpoons$	Cr	– 0,74
$Zn^{2+} + 2e^-$	$\rightleftharpoons$	Zn	– 0,76
$2H_2O + 2e^-$	$\rightleftharpoons$	$H_2(g) + 2OH^-$	– 0,83
$Cr^{2+} + 2e^-$	$\rightleftharpoons$	Cr	– 0,91
$Mn^{2+} + 2e^-$	$\rightleftharpoons$	Mn	– 1,18
$A\ell^{3+} + 3e^-$	$\rightleftharpoons$	$A\ell$	– 1,66
$Mg^{2+} + 2e^-$	$\rightleftharpoons$	Mg	– 2,36
$Na^+ + e^-$	$\rightleftharpoons$	Na	– 2,71
$Ca^{2+} + 2e^-$	$\rightleftharpoons$	Ca	– 2,87
$Sr^{2+} + 2e^-$	$\rightleftharpoons$	Sr	– 2,89
$Ba^{2+} + 2e^-$	$\rightleftharpoons$	Ba	– 2,90
$Cs^+ + e^-$	$\rightleftharpoons$	Cs	- 2,92
$K^+ + e^-$	$\rightleftharpoons$	K	– 2,93
$Li^+ + e^-$	$\rightleftharpoons$	Li	– 3,05

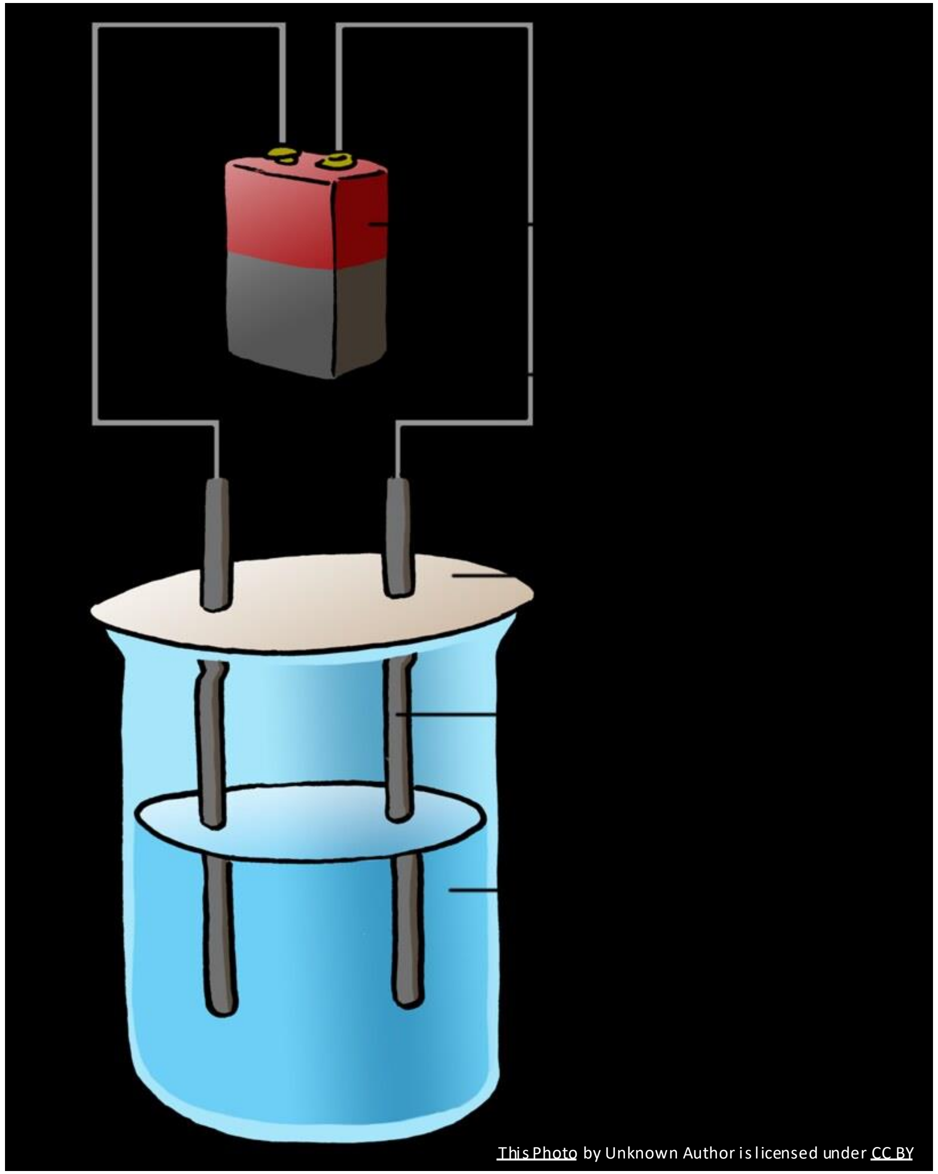
Increasing oxidising ability/*Toenemende oksiderende vermoë*

Increasing reducing ability/*Toenemende reduserende vermoë*

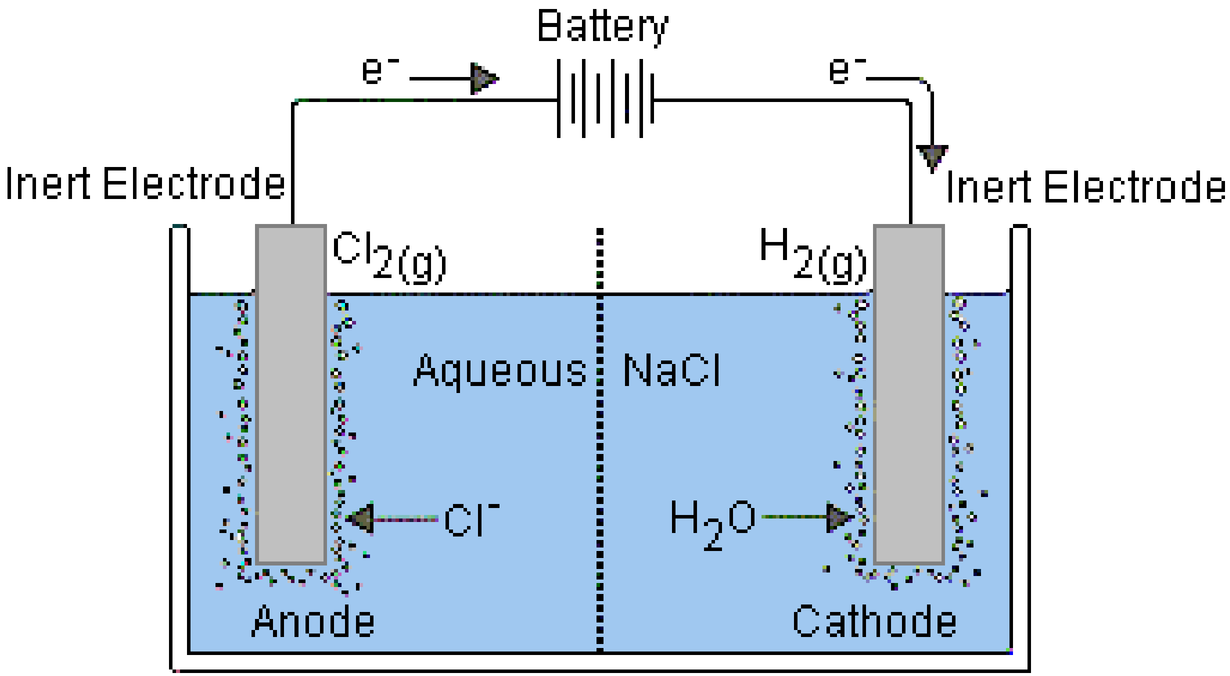
Decomposition of copper chloride

- ANODE
  - Positive electrode
  - Negative Cl^- ions will move to the anode
  - Cl^- ions are OXIDISED to Cl_2
  - Cl^- ions acts as the REDUCING AGENT

- CATHODE
  - Negative electrode
  - Positive Cu^{2+} ions will move to the cathode
  - Cu^{2+} ions are REDUCED to Cu
  - Cu^{2+} ions acts as the OXIDISING AGENT



Electrolysis of a saline solution



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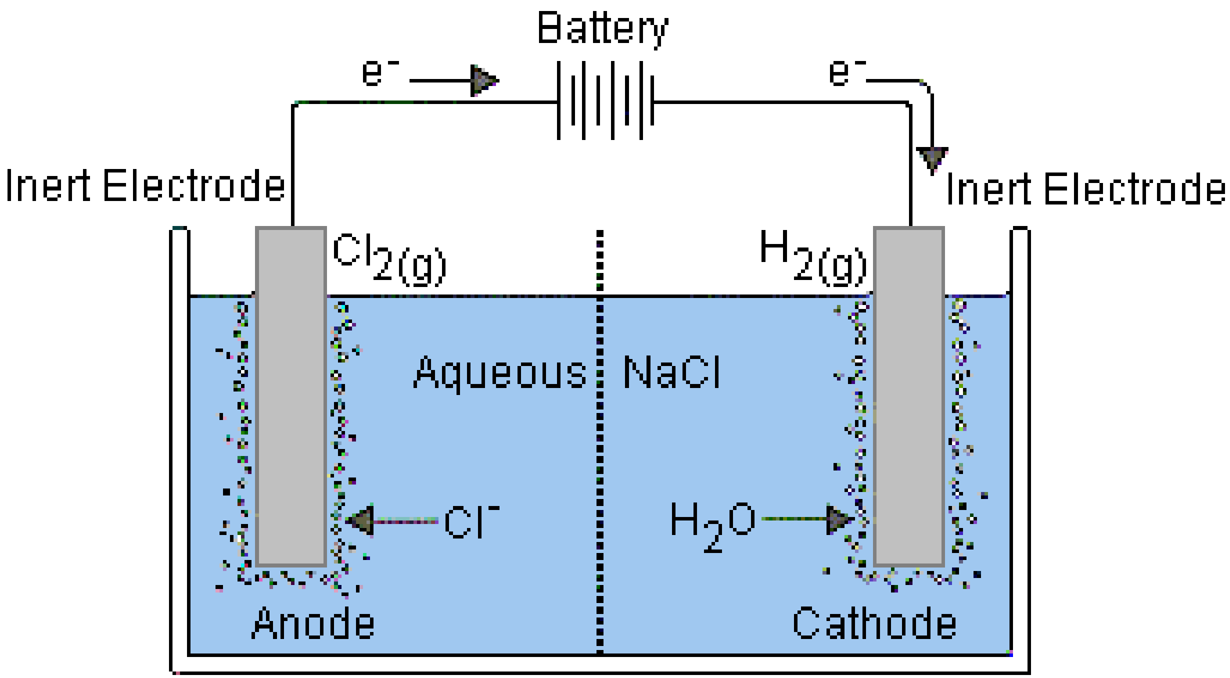
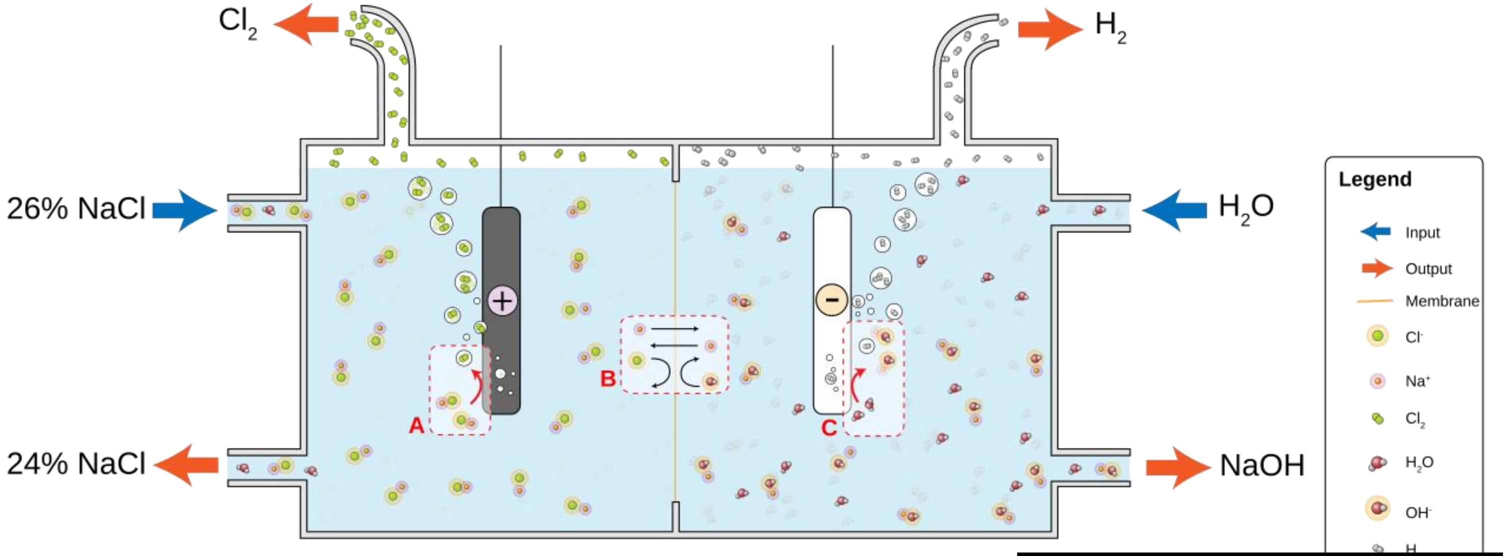
- Ionic salt dissolved in water
- + and - ions of the salt
- H^+ and OH^- of the water
- Metal ions from group I and II elements are weaker oxidizing agents than H_2O
  - Metal ions will remain in solution
  - H_2O will be reduced to form H_2 and OH^-

Electrolysis of a saline solution



- Ionic salt dissolved in water
- + and - ions of the salt
- H^+ and OH^- of the water
- Metal ions that are stronger oxidizing agents than H_2O
  - Cu, Au, Ag, Pt
  - The metal ions will be reduced

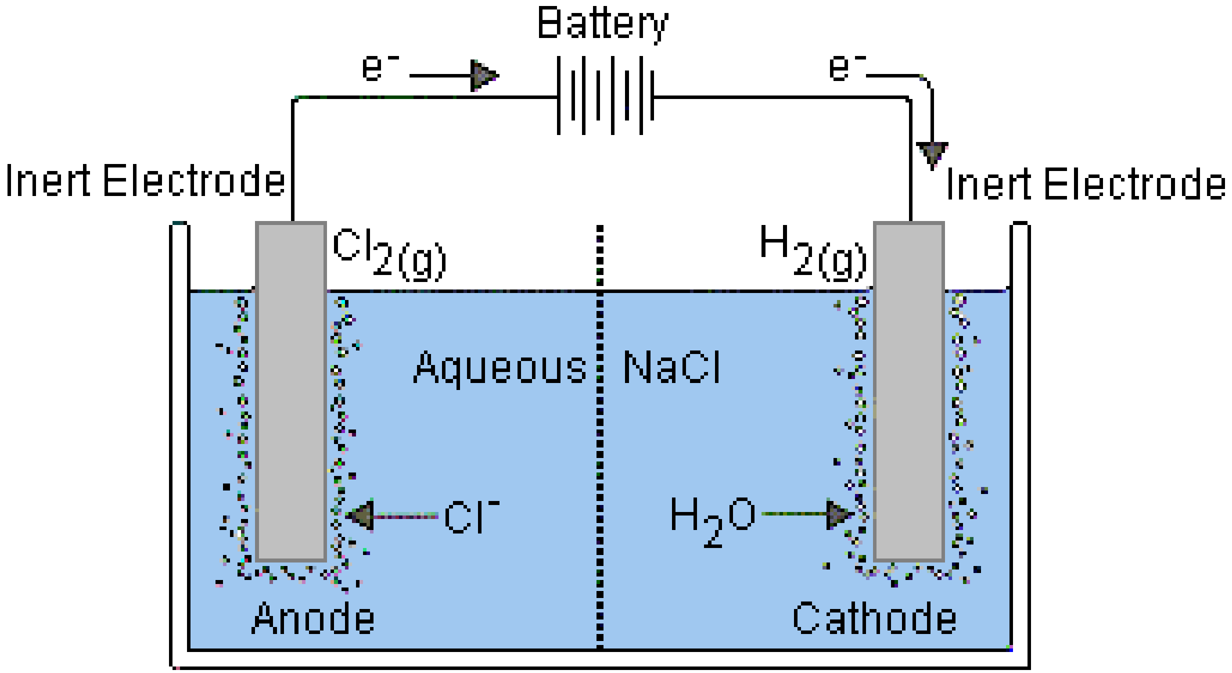
Electrolysis of a saline solution



- EXAMPLE: $CuCl_2$ (aq)
- Oxidation half-reaction: $2Cl^- \rightarrow Cl_2 + 2é$
- Reduction half-reaction: $Cu^{2+} + 2é \rightarrow Cu$

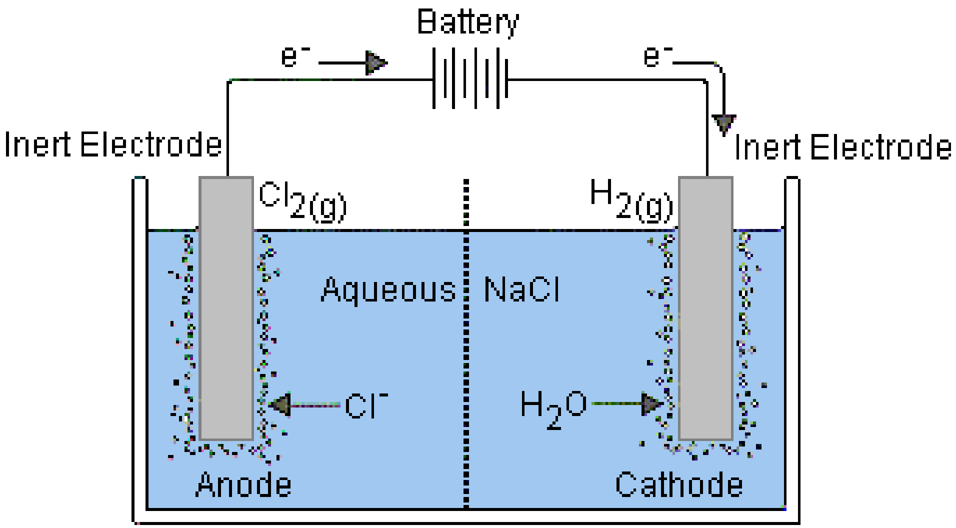
- EXAMPLE: NaCl (aq)
- Oxidation half-reaction: $2Cl^- \rightarrow Cl_2 + 2é$
- Reduction half-reaction: $2H_2O + 2é \rightarrow H_2 + 2OH^-$

Electrolysis of a saline solution



- Ionic salt dissolved in water
- + and - ions of the salt
- H^+ and OH^- of the water
- Negative ions of a halogen are oxidized more easily than H_2O (except F^-)
- Fluorine ion, sulphate, carbonate, nitrate and phosphate are not oxidized, water will be oxidized instead to form O_2 gas.

Electrolysis of a saline solution

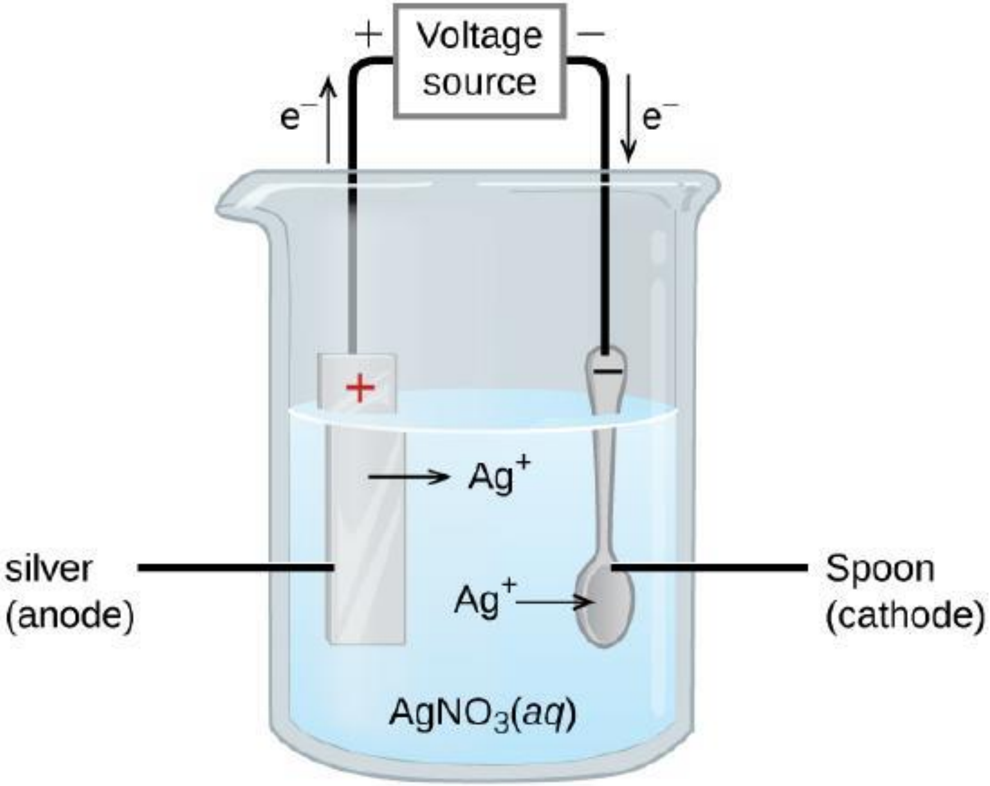


- Negative ions of a halogen are oxidized more easily than H_2O (except F^-)
- Fluorine ion, sulphate, carbonate, nitrate and phosphate are not oxidized, water will be oxidized instead to form O_2 gas.
- EXAMPLE: $NaSO_4$
- Cathode (Reduction half-reaction): $2H_2O + 2é \rightarrow H_2 + 2OH^-$
- Anode (Oxidation half-reaction): $2H_2O \rightarrow O_2 + 4H^+ + 4é$
- Net reaction: $6H_2O \rightarrow 2H_2 + 4OH^- + O_2 + 4H^+$

  $2H_2O \rightarrow 2H_2 + O_2$

Electroplating

- Nickel spoon is coated with Ag
- Object that has to be coated is the cathode
- Anode – metal X (metal that will form the coating)
- Electrolyte – metal salt solution containing metal X-ions
- NB: In electroplating the SAME metal is oxidized and reduced



Electrorefining of copper

- Purification of copper
- Pure copper: Cathode
    - Layer of pure copper forms on electrode
    - Mass of cathode increases
- Impure copper: Anode
    - Impure copper oxidised
    - Mass of anode decreases
- Less reactive elements and compounds from the impure copper will precipitate to the bottom of the container

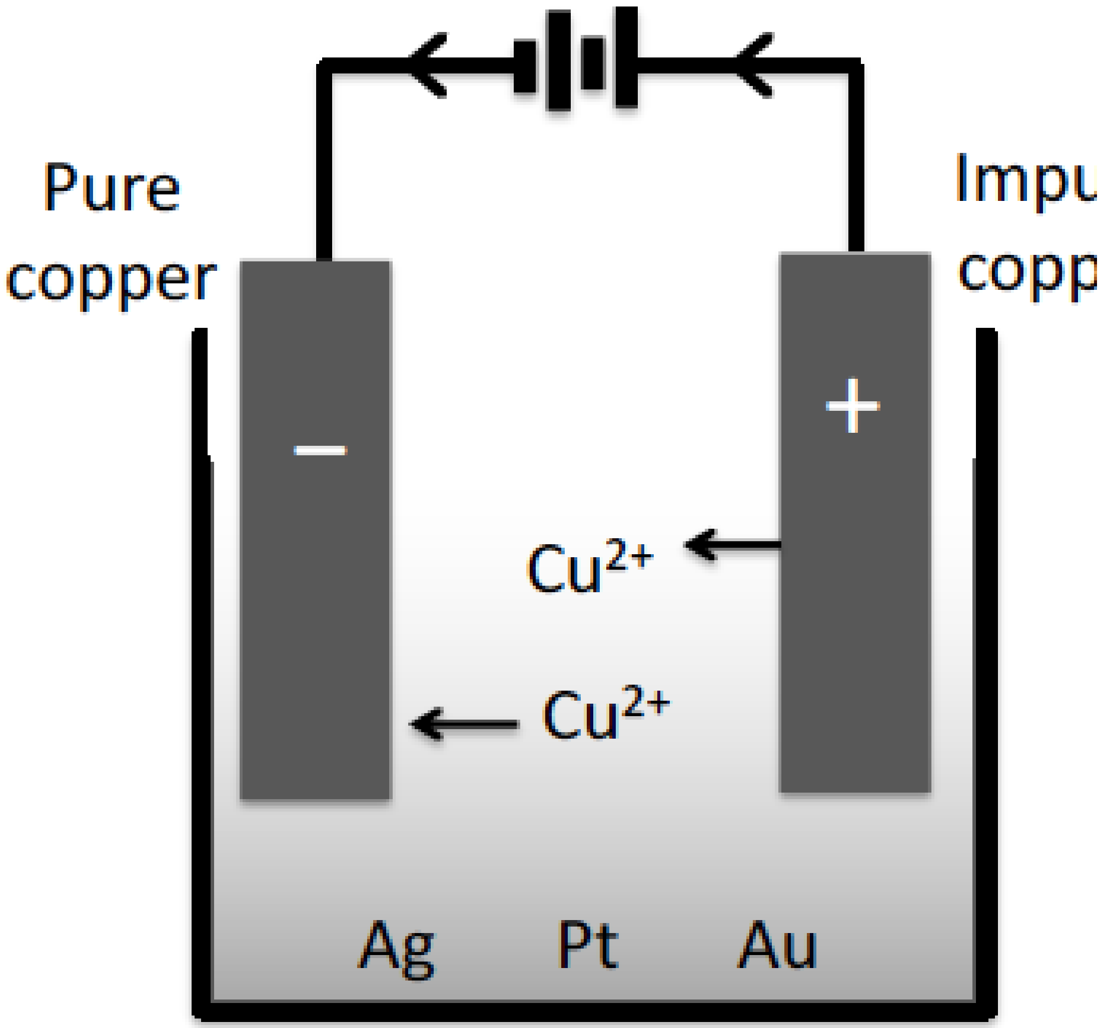
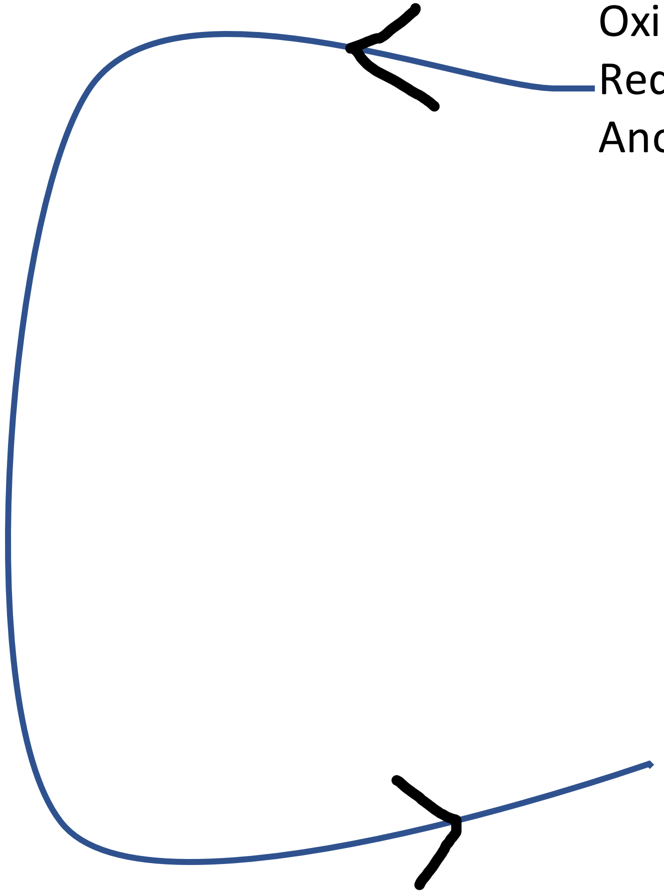


Table 4A



Oxidation half reaction
Reducing agent
Anode (+)

Reduction half reaction
Oxidizing agent
Cathode (-)

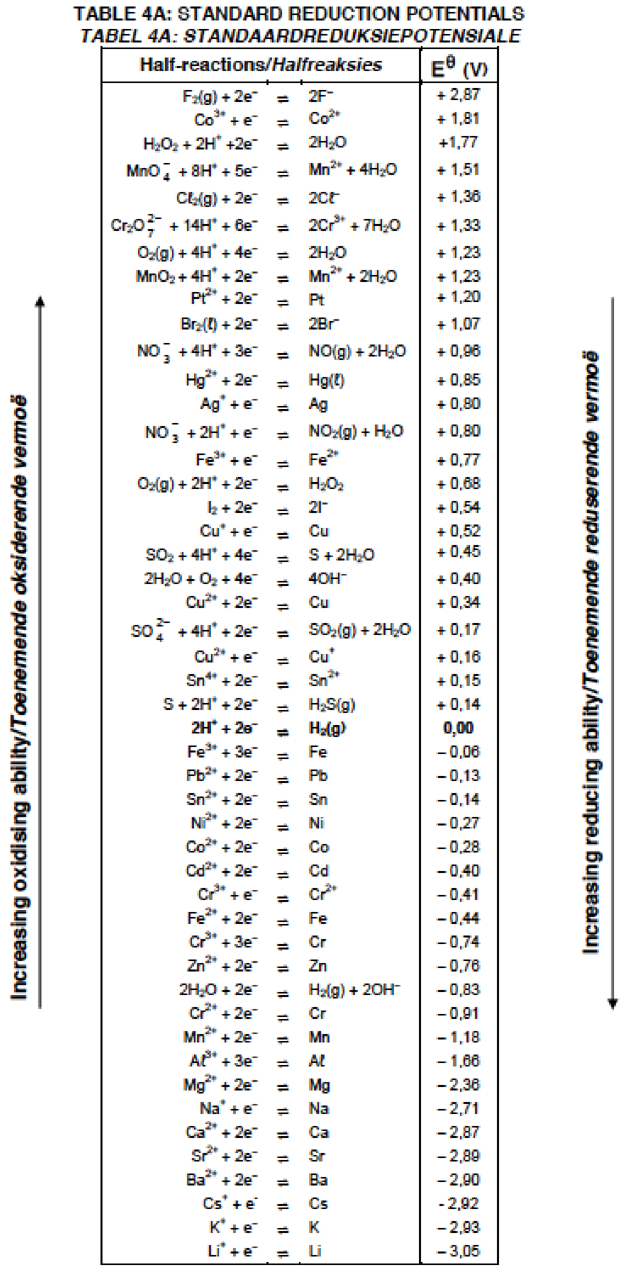
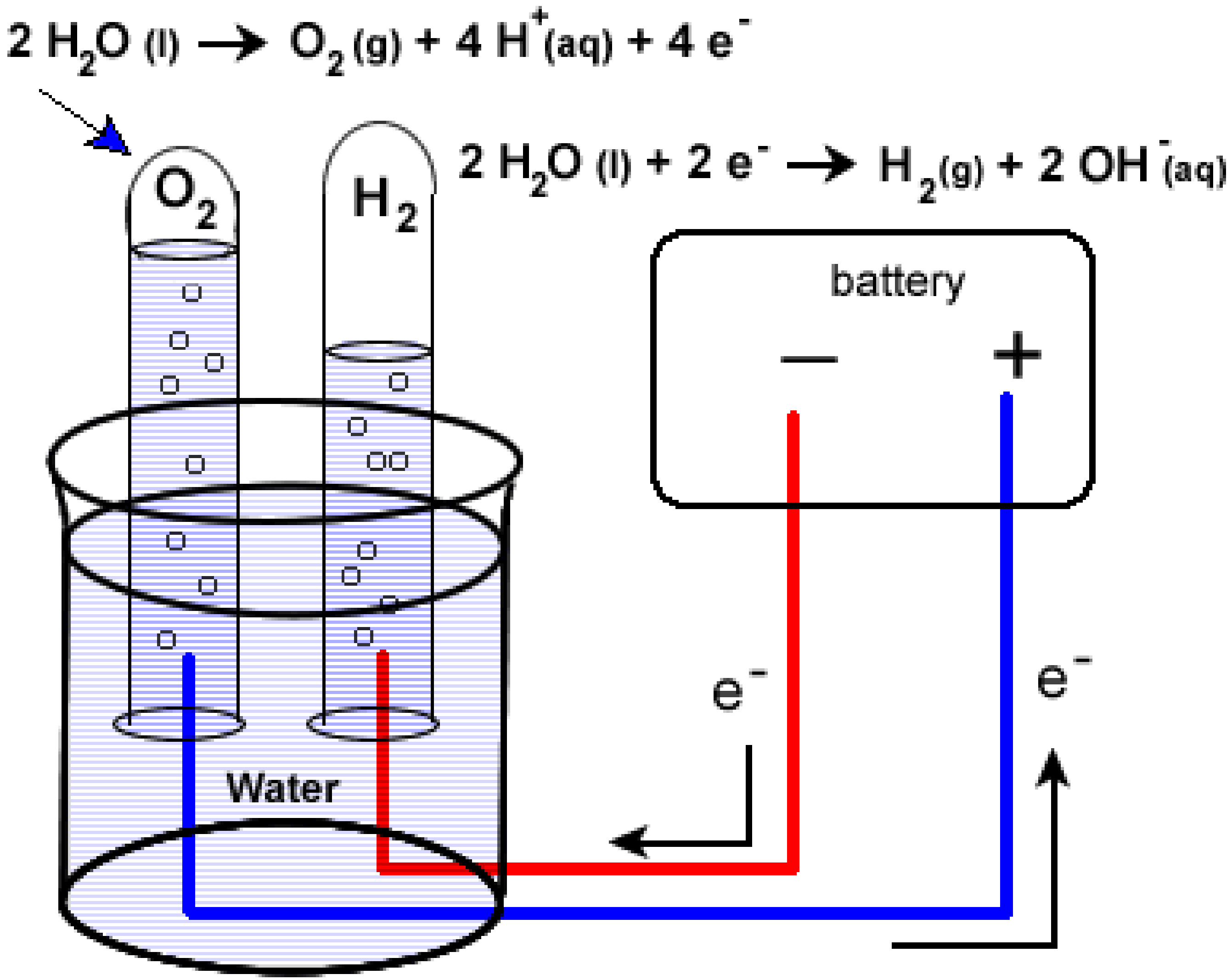


TABLE 4A: STANDARD REDUCTION POTENTIALS
TABEL 4A: STANDAARDREDUKSIEPOTENSIALE

Half-reactions/*Halfreaksies*			**E^{θ} (V)**
$F_2(g) + 2e^-$	$\rightleftharpoons$	$2F^-$	+ 2,87
$Co^{3+} + e^-$	$\rightleftharpoons$	Co^{2+}	+ 1,81
$H_2O_2 + 2H^+ + 2e^-$	$\rightleftharpoons$	$2H_2O$	+1,77
$MnO_4^- + 8H^+ + 5e^-$	$\rightleftharpoons$	$Mn^{2+} + 4H_2O$	+ 1,51
$C\ell_2(g) + 2e^-$	$\rightleftharpoons$	$2C\ell^-$	+ 1,36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightleftharpoons$	$2Cr^{3+} + 7H_2O$	+ 1,33
$O_2(g) + 4H^+ + 4e^-$	$\rightleftharpoons$	$2H_2O$	+ 1,23
$MnO_2 + 4H^+ + 2e^-$	$\rightleftharpoons$	$Mn^{2+} + 2H_2O$	+ 1,23
$Pt^{2+} + 2e^-$	$\rightleftharpoons$	Pt	+ 1,20
$Br_2(\ell) + 2e^-$	$\rightleftharpoons$	$2Br^-$	+ 1,07
$NO_3^- + 4H^+ + 3e^-$	$\rightleftharpoons$	$NO(g) + 2H_2O$	+ 0,96
$Hg^{2+} + 2e^-$	$\rightleftharpoons$	$Hg(\ell)$	+ 0,85
$Ag^+ + e^-$	$\rightleftharpoons$	Ag	+ 0,80
$NO_3^- + 2H^+ + e^-$	$\rightleftharpoons$	$NO_2(g) + H_2O$	+ 0,80
$Fe^{3+} + e^-$	$\rightleftharpoons$	Fe^{2+}	+ 0,77
$O_2(g) + 2H^+ + 2e^-$	$\rightleftharpoons$	H_2O_2	+ 0,68
$I_2 + 2e^-$	$\rightleftharpoons$	$2I^-$	+ 0,54
$Cu^+ + e^-$	$\rightleftharpoons$	Cu	+ 0,52
$SO_2 + 4H^+ + 4e^-$	$\rightleftharpoons$	$S + 2H_2O$	+ 0,45
$2H_2O + O_2 + 4e^-$	$\rightleftharpoons$	$4OH^-$	+ 0,40
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	Cu	+ 0,34
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2(g) + 2H_2O$	+ 0,17
$Cu^{2+} + e^-$	$\rightleftharpoons$	Cu^+	+ 0,16
$Sn^{4+} + 2e^-$	$\rightleftharpoons$	Sn^{2+}	+ 0,15
$S + 2H^+ + 2e^-$	$\rightleftharpoons$	$H_2S(g)$	+ 0,14
$2H^+ + 2e^-$	$\rightleftharpoons$	**$H_2(g)$**	**0,00**
$Fe^{3+} + 3e^-$	$\rightleftharpoons$	Fe	− 0,06
$Pb^{2+} + 2e^-$	$\rightleftharpoons$	Pb	− 0,13
$Sn^{2+} + 2e^-$	$\rightleftharpoons$	Sn	− 0,14
$Ni^{2+} + 2e^-$	$\rightleftharpoons$	Ni	− 0,27
$Co^{2+} + 2e^-$	$\rightleftharpoons$	Co	− 0,28
$Cd^{2+} + 2e^-$	$\rightleftharpoons$	Cd	− 0,40
$Cr^{3+} + e^-$	$\rightleftharpoons$	Cr^{2+}	− 0,41
$Fe^{2+} + 2e^-$	$\rightleftharpoons$	Fe	− 0,44
$Cr^{3+} + 3e^-$	$\rightleftharpoons$	Cr	− 0,74
$Zn^{2+} + 2e^-$	$\rightleftharpoons$	Zn	− 0,76
$2H_2O + 2e^-$	$\rightleftharpoons$	$H_2(g) + 2OH^-$	− 0,83
$Cr^{2+} + 2e^-$	$\rightleftharpoons$	Cr	− 0,91
$Mn^{2+} + 2e^-$	$\rightleftharpoons$	Mn	− 1,18
$A\ell^{3+} + 3e^-$	$\rightleftharpoons$	$A\ell$	− 1,66
$Mg^{2+} + 2e^-$	$\rightleftharpoons$	Mg	− 2,36
$Na^+ + e^-$	$\rightleftharpoons$	Na	− 2,71
$Ca^{2+} + 2e^-$	$\rightleftharpoons$	Ca	− 2,87
$Sr^{2+} + 2e^-$	$\rightleftharpoons$	Sr	− 2,89
$Ba^{2+} + 2e^-$	$\rightleftharpoons$	Ba	− 2,90
$Cs^+ + e^-$	$\rightleftharpoons$	Cs	- 2,92
$K^+ + e^-$	$\rightleftharpoons$	K	− 2,93
$Li^+ + e^-$	$\rightleftharpoons$	Li	− 3,05

Increasing oxidising ability/*Toenemende oksiderende vermoë*

Increasing reducing ability/*Toenemende reduserende vermoë*



$$2\ H_2O\ (l) \rightarrow O_2\ (g) + 4\ H^+(aq) + 4\ e^-$$

$$2\ H_2O\ (l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$$

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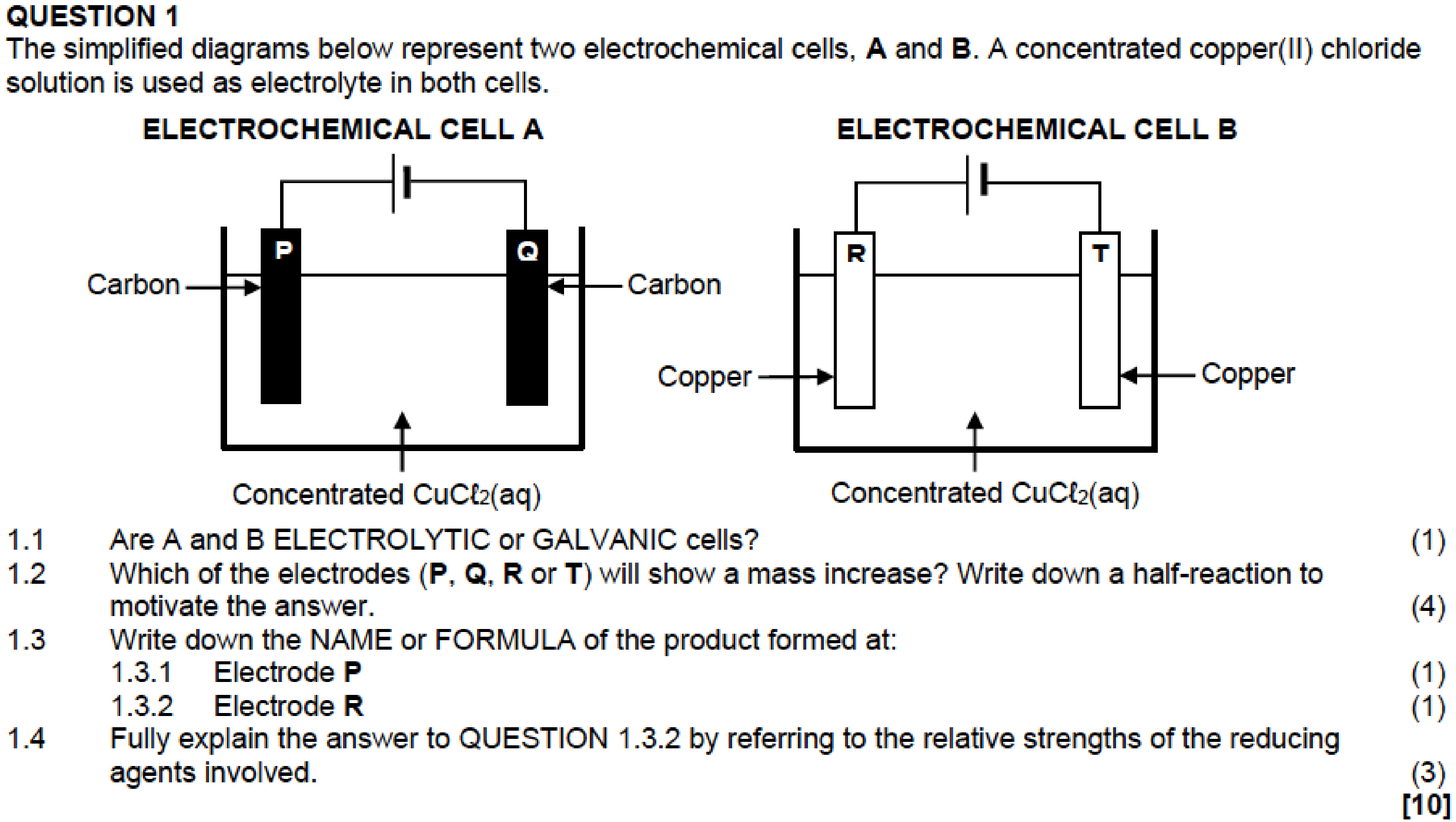
Example 1



QUESTION 1

The simplified diagrams below represent two electrochemical cells, **A** and **B**. A concentrated copper(II) chloride solution is used as electrolyte in both cells.

1.1 Are A and B ELECTROLYTIC or GALVANIC cells? (1)

1.2 Which of the electrodes (**P**, **Q**, **R** or **T**) will show a mass increase? Write down a half-reaction to motivate the answer. (4)

1.3 Write down the NAME or FORMULA of the product formed at:

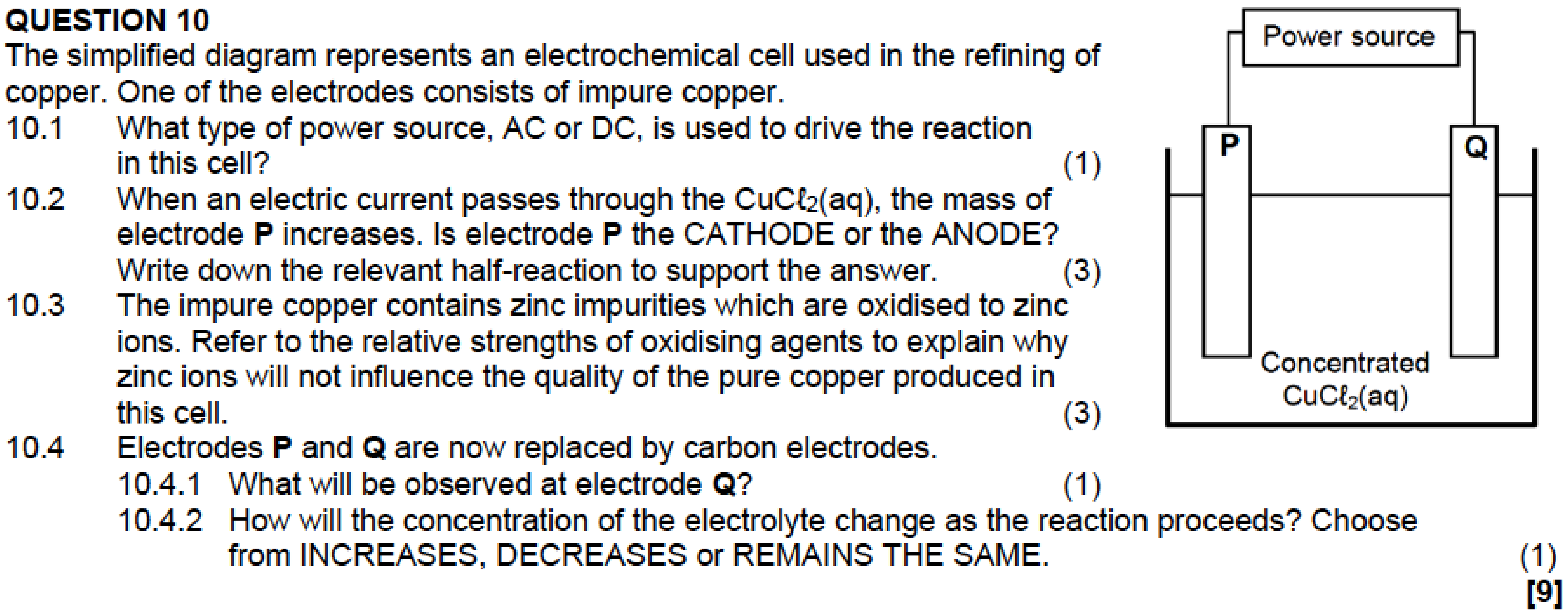
1.3.1 Electrode **P** (1)

1.3.2 Electrode **R** (1)

1.4 Fully explain the answer to QUESTION 1.3.2 by referring to the relative strengths of the reducing agents involved. (3)

[10]

Example 2



QUESTION 10

The simplified diagram represents an electrochemical cell used in the refining of copper. One of the electrodes consists of impure copper.

10.1 What type of power source, AC or DC, is used to drive the reaction in this cell? (1)

10.2 When an electric current passes through the $CuC\ell_2$(aq), the mass of electrode **P** increases. Is electrode **P** the CATHODE or the ANODE? Write down the relevant half-reaction to support the answer. (3)

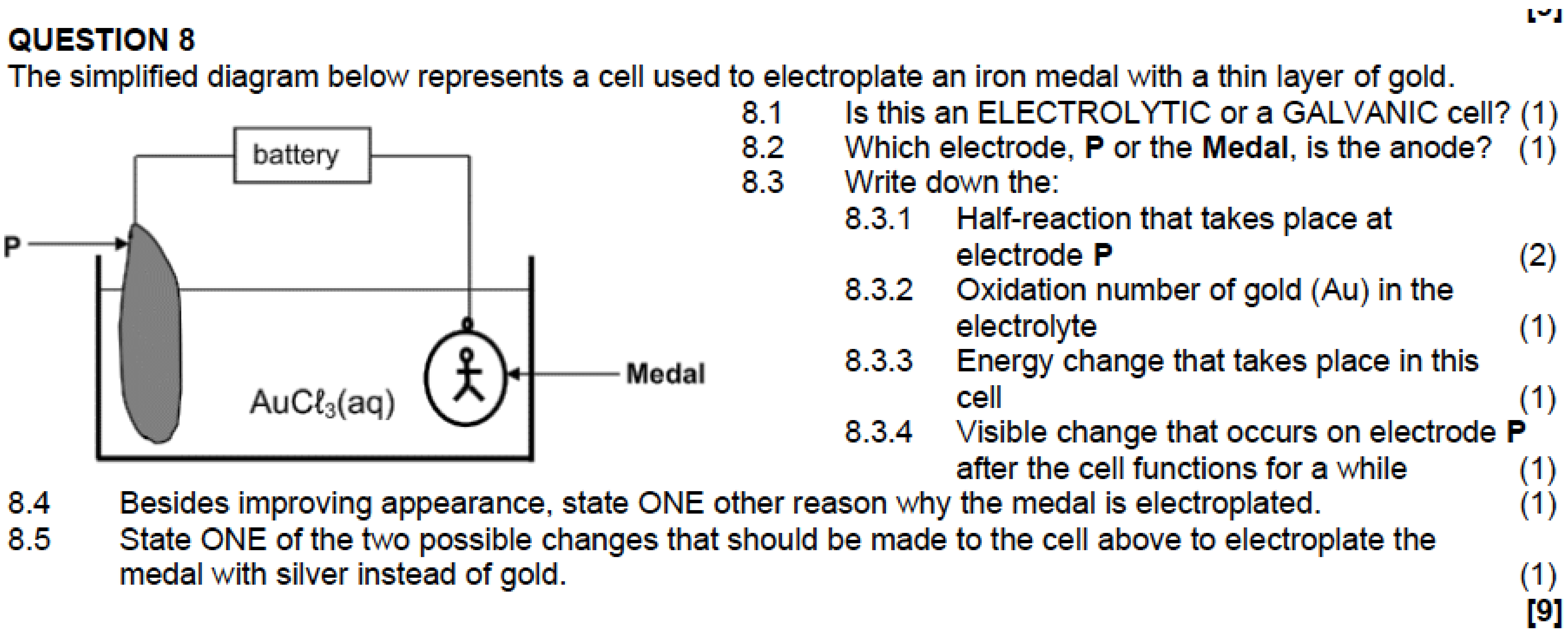
10.3 The impure copper contains zinc impurities which are oxidised to zinc ions. Refer to the relative strengths of oxidising agents to explain why zinc ions will not influence the quality of the pure copper produced in this cell. (3)

10.4 Electrodes **P** and **Q** are now replaced by carbon electrodes.

- 10.4.1 What will be observed at electrode **Q**? (1)
- 10.4.2 How will the concentration of the electrolyte change as the reaction proceeds? Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)

[9]

Example 3



QUESTION 8

The simplified diagram below represents a cell used to electroplate an iron medal with a thin layer of gold.

8.1 Is this an ELECTROLYTIC or a GALVANIC cell? (1)

8.2 Which electrode, **P** or the **Medal**, is the anode? (1)

8.3 Write down the:

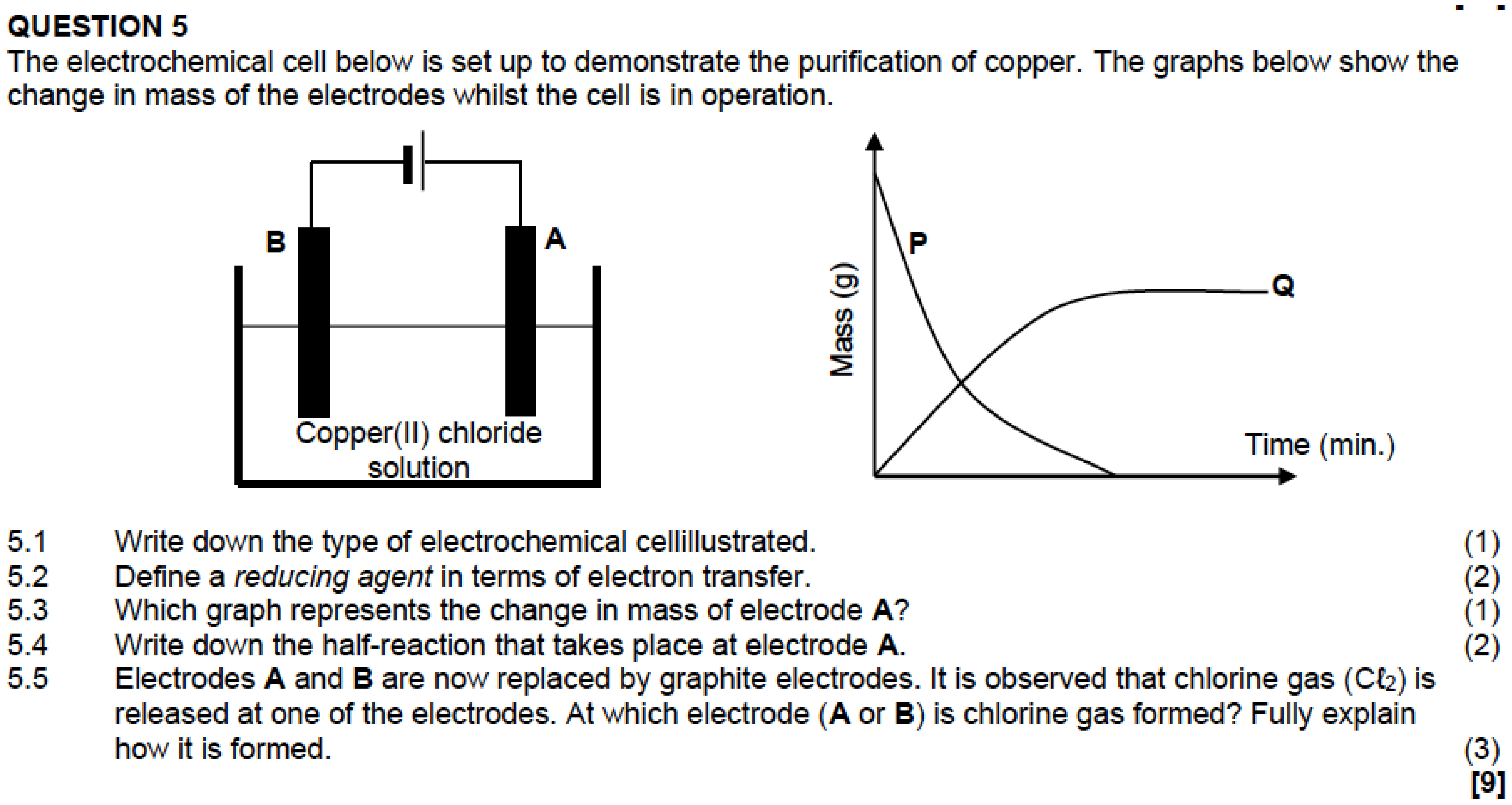
- 8.3.1 Half-reaction that takes place at electrode **P** (2)
- 8.3.2 Oxidation number of gold (Au) in the electrolyte (1)
- 8.3.3 Energy change that takes place in this cell (1)
- 8.3.4 Visible change that occurs on electrode **P** after the cell functions for a while (1)

8.4 Besides improving appearance, state ONE other reason why the medal is electroplated. (1)

8.5 State ONE of the two possible changes that should be made to the cell above to electroplate the medal with silver instead of gold. (1)

[9]

Example 4



QUESTION 5

The electrochemical cell below is set up to demonstrate the purification of copper. The graphs below show the change in mass of the electrodes whilst the cell is in operation.

5.1 Write down the type of electrochemical cellillustrated. (1)

5.2 Define a *reducing agent* in terms of electron transfer. (2)

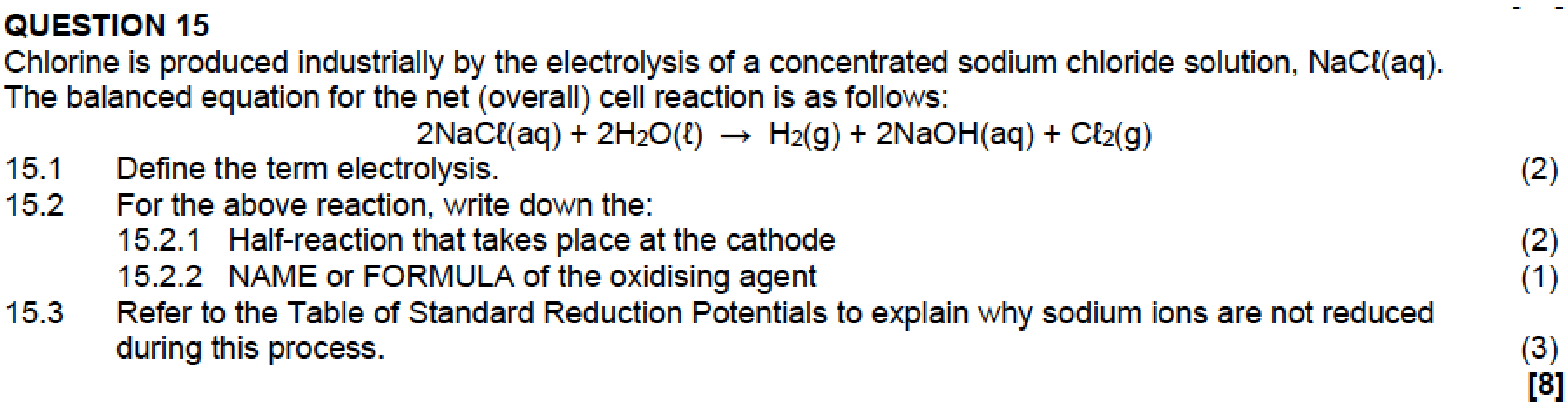
5.3 Which graph represents the change in mass of electrode **A**? (1)

5.4 Write down the half-reaction that takes place at electrode **A**. (2)

5.5 Electrodes **A** and **B** are now replaced by graphite electrodes. It is observed that chlorine gas ($C\ell_2$) is released at one of the electrodes. At which electrode (**A** or **B**) is chlorine gas formed? Fully explain how it is formed. (3)

[9]

Example 5



QUESTION 15

Chlorine is produced industrially by the electrolysis of a concentrated sodium chloride solution, NaCℓ(aq). The balanced equation for the net (overall) cell reaction is as follows:

$$2NaC\ell(aq) + 2H_2O(\ell) \rightarrow H_2(g) + 2NaOH(aq) + C\ell_2(g)$$

15.1 Define the term electrolysis. (2)

15.2 For the above reaction, write down the:

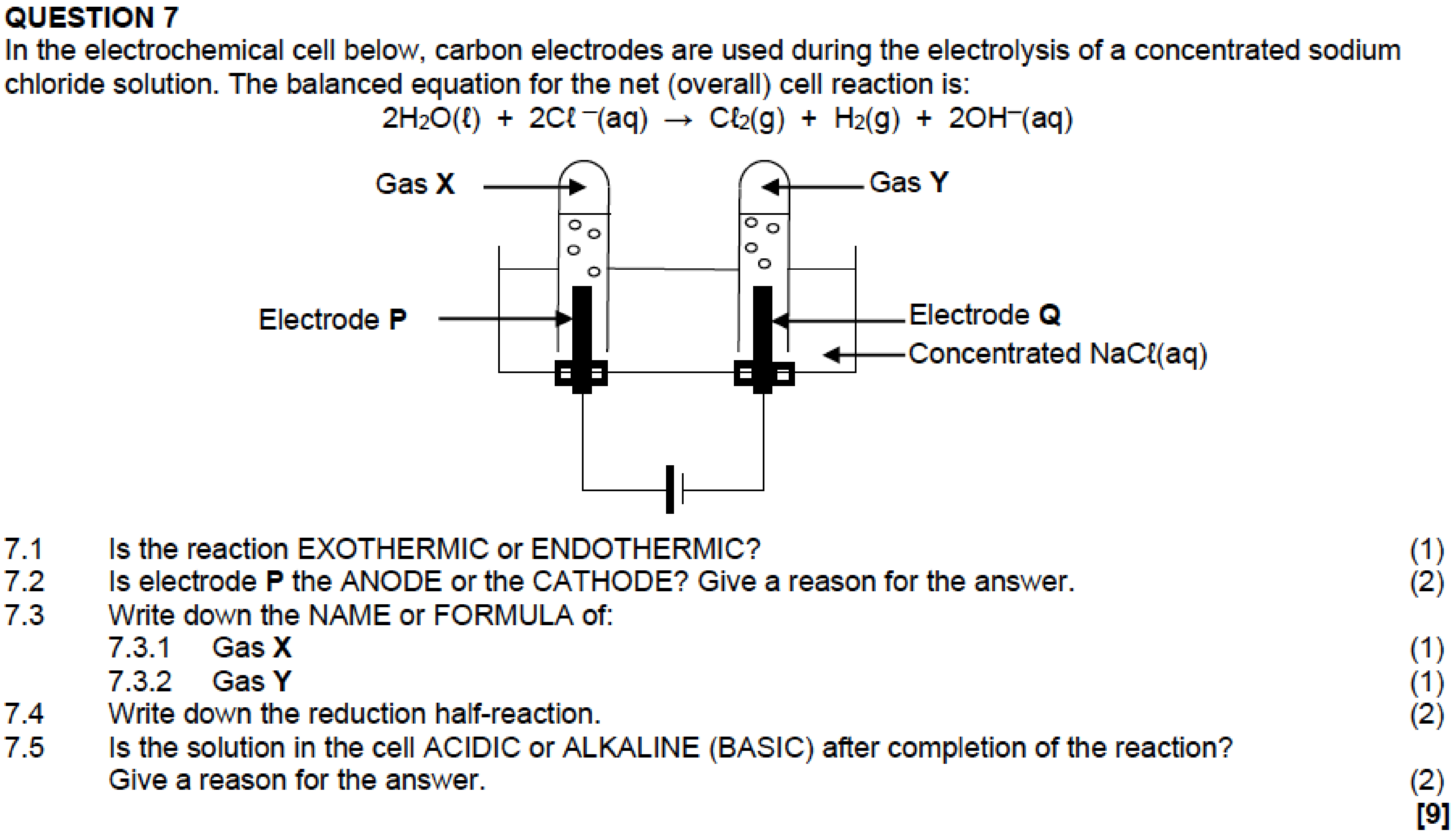
  15.2.1 Half-reaction that takes place at the cathode (2)

  15.2.2 NAME or FORMULA of the oxidising agent (1)

15.3 Refer to the Table of Standard Reduction Potentials to explain why sodium ions are not reduced during this process. (3)

[8]

Example 6



QUESTION 7

In the electrochemical cell below, carbon electrodes are used during the electrolysis of a concentrated sodium chloride solution. The balanced equation for the net (overall) cell reaction is:

$$2H_2O(\ell) + 2C\ell^-(aq) \rightarrow C\ell_2(g) + H_2(g) + 2OH^-(aq)$$

7.1 Is the reaction EXOTHERMIC or ENDOTHERMIC? (1)
7.2 Is electrode **P** the ANODE or the CATHODE? Give a reason for the answer. (2)
7.3 Write down the NAME or FORMULA of:
  7.3.1 Gas **X** (1)
  7.3.2 Gas **Y** (1)
7.4 Write down the reduction half-reaction. (2)
7.5 Is the solution in the cell ACIDIC or ALKALINE (BASIC) after completion of the reaction? Give a reason for the answer. (2)
[9]