Iron is more reactive than copper.

(i) Draw a labelled diagram of a voltaic cell made from an Fe(s) / Fe2+(aq) half-cell connected to a Cu(s) / Cu2+(aq) half-cell. In your diagram identify the positive electrode (cathode), the negative electrode (anode) and the direction of electron flow in the external circuit.

(ii) Deduce the half-equations for the reactions taking place at the positive electrode (cathode) and negative electrode (anode) of this voltaic cell.

 (iii) Deduce the overall equation for the reaction taking place in the voltaic cell and determine which species acts as the oxidizing agent and which species has been reduced.

(i)

correct diagram including voltmeter/meter, 4 correct species (state
symbols not required) and connecting wires;
*No credit if wires to electrodes immersed in the solutions.*

 labelled salt bridge;
*Do not accept name of salt (e.g. potassium nitrate) in place
of salt bridge.*

 correctly labelled electrodes (+)/cathode and (–)/anode;
flow of electrons from Fe to Cu in external circuit; 4

(ii) positive electrode: Cu2+ + 2e– → Cu;
negative electrode: Fe → Fe2+ + 2e–; 2
*Award* ***[1]*** *if equations correct but at wrong electrodes or
if electrodes are missing.
Award* ***[2]*** *for correct equations if electrodes are missing but
were correctly labelled in diagram.
Accept e instead of e–.
Ignore state symbols.
Penalize  once only in equations in (ii) and (iii).*

(iii) Fe + Cu2+ → Fe2+ + Cu;
*Ignore state symbols.*

 Cu2+ is the oxidizing agent **and** the species that is reduced

1.

(i) Draw an annotated diagram of a voltaic cell composed of a magnesium electrode in 1.0 mol dm–3 magnesium nitrate solution and a silver electrode in 1.0 mol dm–3 silver nitrate solution. State the direction of electron flow on your diagram.

(ii) Deduce half-equations for the oxidation and reduction reactions

(i)

correctly labelled electrodes **and** solutions;
labelled salt bridge;
voltmeter;
*Allow bulb or ammeter.*

 direction of electron flow; 4

(ii) *Oxidation:*Mg(s) → Mg2+(aq) + 2e–;

 Reduction:
Ag+(aq) + e– → Ag(s);

1.

Consider the following three redox reactions.

 Cd(s) + Ni2+(aq) → Cd2+(aq) + Ni(s)

 Ni(s) + 2Ag+(aq) → Ni2+(aq) + 2Ag(s)

 Zn(s) + Cd2+(aq) → Zn2+(aq) + Cd(s)

(i) Deduce the order of reactivity of the four metals, cadmium, nickel, silver and zinc and list in order of **decreasing** reactivity.

(2)

(ii) Identify the best oxidizing agent and the best reducing agent.

(i) Zn > Cd > Ni > Ag
Zn most reactive;
rest of order correct; 2

(ii) *Best oxidizing agent:*Ag+;
*Do not accept Ag.*

 *Best reducing agent:*Zn;
*Do not accept Zn2+.*

1.

(i) Solid sodium chloride does not conduct electricity but molten sodium chloride does.
Explain this difference.

**(2)**

(ii) Outline what happens in an electrolytic cell during the electrolysis of molten sodium chloride using inert electrodes. Deduce equations for the reactions occurring at each electrode.

(i) sodium chloride crystals consist of ions in a (rigid) lattice / ions cannot
move (to electrodes) / *OWTTE*;
when melted ions free to move / ions move when potential difference/
voltage applied; 2

(ii) positive sodium ions/Na+ move to negative electrode/cathode **and** negative
chloride ions/Cl– move to positive electrode/anode;

 electrons released to positive electrode/anode by negative ions and accepted
from negative electrode/cathode by positive ions / reduction occurs at the
negative electrode/cathode **and** oxidation occurs at the positive electrode/
anode / Na+ ions are reduced **and** Cl– ions are oxidized;

 *(Positive electrode/anode):*2Cl– → Cl2 + 2e– / Cl– → Cl2 + e–;

 *(Negative electrode/cathode):*2Na+ + 2e– → 2Na / Na+ + e– → Na;

1.

The word *redox* comes from a combination of the terms *reduction* and *oxidation*.
Redox reactions affect our daily lives.

 The overall reaction that takes place in a voltaic cell is shown below.

 Pb(s) + PbO2(s) + 2H2SO4(aq) → 2PbSO4(s) + 2H2O(l)

(i) Determine the oxidation number of lead in Pb, PbO2 and PbSO4.

**(1)**

(ii) Deduce the oxidation and reduction half-equations taking place at the negative lead electrode (anode) and the positive lead(IV) oxide electrode (cathode). Deduce the oxidizing and reducing agents and state the direction of the electron flow between the electrodes.

**(4)**

(iii) In order to determine the position of three metals in a reactivity series, the metals were placed in different solutions of metal ions. The table below summarizes whether or not a reaction occurred.

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Ag+(aq)** | **Cu2+(aq)** | **Pb2+(aq)** |
| **Ag(s)** |  | No reaction | No reaction |
| **Cu(s)** | Reaction |  | No reaction |
| **Pb(s)** | Reaction | Reaction |  |

 State the equations for the **three** reactions that take place. Use this information to place the metals Ag, Cu and Pb in a reactivity series, with the strongest reducing agent first, and explain your reasoning.

(i) Pb: 0, PbO2: +4, PbSO4: +2;
*Need sign for mark.
Do not accept notations such as 4+, 2+ or IV, II.* 1

(ii) *Negative/–/anode*Pb(s) + SO42–(aq) → PbSO4(s) + 2e– / Pb(s) → Pb2+(aq) + 2e–;

 *Positive/+/cathode*PbO2(s) + 4H+(aq) + SO42–(aq) + 2e– → PbSO4(s) + 2H2O(l) /
PbO2(s) + 4H+(aq) + 2e– → Pb2+(aq) + 2H2O(l) /
PbO2(s) + H2SO4(aq) + 2H+(aq) + 2e– → PbSO4(s) + 2H2O(l);
*Accept Pb4+ + 2e– → Pb2+*.
*Ignore state symbols.
Allow e instead of e–.*

 oxidizing agent is PbO2 / lead(IV) oxide/lead dioxide **and** reducing agent
is Pb/lead;
from negative/–/anode/Pb to positive/+/cathode/PbO2 (through the external
circuit/wire); 4

(iii) Pb(s) + Cu2+(aq) → Pb2+(aq) + Cu(s)
Pb(s) + 2Ag+(aq) → Pb2+(aq) + 2Ag(s)
Cu(s) + 2Ag+(aq) → Cu2+(aq) + 2Ag(s)
*Award* ***[2]*** *for three correct, award* ***[1]*** *for any two correct, one correct
scores no mark.
Ignore state symbols.
Penalize unbalanced equations once only.*

 Pb is a stronger reducing agent than Cu and/or Ag / Pb most reactive as it
can reduce/displace both Cu2+ and Ag+;
Cu is a stronger reducing agent than Ag but not Pb / Cu in the middle (of the
three) as it can reduce/displace Ag+ but not Pb2+;

1.

Fertilizers may cause health problems for babies because nitrates can change into nitrites in water used for drinking.

(i) Define *oxidation* in terms of oxidation numbers.

(1)

(ii) Deduce the oxidation states of nitrogen in the nitrate, NO3–, and nitrite, NO2–, ions.

(i) increase in the oxidation number; 1

(ii) (NO3)+ 5 **and** (NO2–) + 3;
Accept V and III.

1.

Nitric acid reacts with silver in a redox reaction.

 \_\_ Ag(s) + \_\_ NO3–(aq) + \_\_\_\_ → \_\_ Ag+(aq) + \_\_ NO(g) + \_\_\_\_

 Using oxidation numbers, deduce the complete balanced equation for the reaction showing all the reactants and products.

change in oxidation numbers: Ag from 0 to +1 **and** N from +5 to +2;
*Do not penalize missing charges on numbers.*

 balanced equation: 3Ag + NO3– + 4H+ → 3Ag+ + NO + 2H2O

1.

(i) Define *oxidation* in terms of oxidation numbers.

**(1)**

(ii) Describe using a labelled diagram, the essential components of an electrolytic cell.

**(3)**

(iii) Explain why solid sodium chloride does not conduct electricity but **molten** sodium chloride does.

**(2)**

(iv) Molten sodium chloride undergoes electrolysis in an electrolytic cell. For each electrode deduce the half-equation and state whether oxidation or reduction takes place. Deduce the equation of the overall cell reaction including state symbols.

**(5)**

(v) Electrolysis has made it possible to obtain reactive metals such as aluminium from their ores, which has resulted in significant developments in engineering and technology. State **one** reason why aluminium is preferred to iron in many uses.

**(1)**

(vi) Outline **two** differences between an electrolytic cell and a voltaic cell.

**(2)**

(i) increase in the oxidation number; 1

(ii) *Annotated diagram of cell showing:*power supply/battery;
electrolyte;
cathode/negative electrode **and** anode/positive electrode;
 3

(iii) (solid) ions in a lattice / ions cannot move;
(molten) ions mobile / ions free to move; 2

(iv) reduction occurs at the cathode/negative electrode **and** oxidation occurs
at the anode/positive electrode;
*Cathode/negative electrode:* Na+ + e– → Na;
*Anode/positive electrode:* 2Cl– → Cl2 + 2e– / Cl– → Cl2 + e–;
*Award* ***[1 max]*** *if the two electrodes are not labelled/labelled incorrectly
for the two half-equations.*

 *Overall cell reaction:* Na+(l) + Cl–(l) → Na(l) + Cl2(g)
*Award* ***[1]*** *for correct equation and* ***[1]*** *for correct state symbols.
Allow NaCl(l) instead of Na+(l) and Cl–(l).* 5

(v) Al does not corrode/rust / Al is less dense/better conductor/more
malleable;
*Accept Al is a lighter (metal compared to Fe).
Accept converse argument.* 1

(vi) electrolytic cell converts electrical energy to chemical energy **and** voltaic
cell converts chemical energy to electrical energy / electrolytic cell uses
electricity to carry out a (redox) chemical reaction **and** voltaic cell uses a
(redox) chemical reaction to produce electricity / electrolytic cell requires
a power supply **and** voltaic cell does not;

 electrolytic cell involves a non-spontaneous (redox) reaction **and** voltaic
cell involves a spontaneous (redox) reaction;

 in an electrolytic cell, cathode is negative and anode is positive **and** *vice-versa*
for a voltaic cell / electrolytic cell, anode is positive and voltaic cell,
anode is negative / electrolytic cell, cathode is negative and voltaic
cell, cathode is positive;

 voltaic cell has two separate solutions **and** electrolytic cell has one
solution / voltaic cell has salt bridge and electrolytic cell has no salt bridge;

 electrolytic cell, oxidation occurs at the positive electrode/anode **and** voltaic
cell, oxidation occurs at the negative electrode/anode and *vice-versa*; 2 max