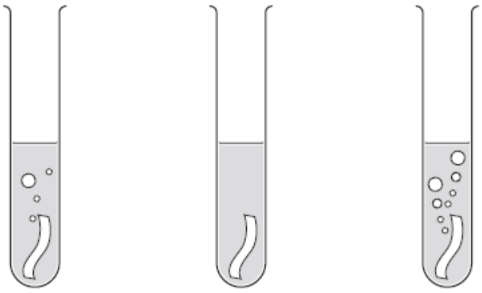
A. Reactivity of metals – The reactivity series, metal oxides and extractions

1. Three metals, X, Y and Z were put into water. The reactions are shown below:



a) Use the diagrams to put metals X, Y and Z in order of reactivity, starting with the most reactive. (1)
Z X Y (1)

b) When a metal reacts with water, it produces hydrogen gas and a metal hydroxide. Describe how you can test for the products. (2)
Hydrogen – gives a squeak pop with a lit splint (1)
Metal hydroxide – turns blue with litmus/purple with universal indicator (1)

c) Give two variables that should be controlled in this investigation. (2)
any two from:
- ***same temperature of the water***
- ***same mass / number of moles of the metal***
- ***same surface area of the metal***

2. A piece of magnesium ribbon was added to dilute hydrochloric acid.

   a) Give two observations that are evidence for a chemical reaction taking place. (2)
   Any two from: gas given off / fizzing, magnesium gets smaller then disappears / thermal energy released / surroundings get warmer/ temperature increases (2)

   b) Write the word and balanced symbol equation, including state symbols, for the reaction. (4)
   magnesium + sulfuric acid → magnesium sulfate + hydrogen [1]
   $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$ [1 – formula equation/1 – balancing/1 – state symbols]

3. The reaction between aluminium powder and iron(III) oxide (Fe_2O_3) is used in the rail industry.

   a) Write a word equation and balanced symbol equation for the reaction that takes place. (3)
   Aluminium + iron(III) oxide → aluminium oxide + iron (1)
   $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$ (2)

   b) Compare the reaction above to the reaction with powdered aluminium and copper(II) oxide and explain why there is a difference. (1)
   Aluminium reacts more vigorously with copper(II) oxide than with iron(III) oxide. (1)
   Copper is lower than iron in the reactivity series, so the reaction is more violent. (1)

4. A student carried out some displacement reactions using three metals and three sulfate solutions. The results are shown in the table below:

	Iron sulfate ($FeSO_4$)	Copper sulfate ($CuSO_4$)	Magnesium sulfate ($MgSO_4$)
Iron (Fe)		✓	×
Copper (Cu)	×		×
Magnesium (Mg)	✓	✓	

a) i) Explain what is observed when iron reacts with copper sulfate. (2)
iron nail turns (from grey to) brown (1)
solution turns (from blue to) pale green (1)

ii) HT: Write an ionic equation for the reaction between iron and copper sulfate solution. (2)
$Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$ (1 mark for reactants and 1 mark for products)

b) Explain why there is no observation between copper and iron sulfate. (2)
Copper does not react with iron sulfate (1) (because) copper is less reactive than iron (1)

c) i) Explain what is observed when magnesium reacts with iron sulfate. (2)
magnesium ribbon turns (from silver to) grey (1)
solution turns (from colourless to) pale green (1)

ii) HT: Write a half equation to show the reduction of iron ions (Fe^{2+}) when magnesium reacts with iron sulfate. Use the half equation to explain why Fe^{2+} ions are reduced.
$Fe^{2+} + 2e^- \rightarrow Fe$ (1)
(Fe^{2+} ions are reduced because they have) gained electrons (1))

B. Reactions of metals part 1 – Metals & acids and strong & weak acids (HT)

1. Zinc reacts with hydrochloric acid.
a) Write a word and a balanced symbol equation, with state symbols, to show this reaction. (3)
zinc + hydrochloric acid → zinc chloride + hydrogen (1)
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ [2]

b) HT: write an ionic equation for the reaction. (2)
$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$ [2]

c) HT: Give both half equations to show the electron transfers taking place. (2)
$Zn \rightarrow Zn^{2+} + 2e^-$ [1] $2H^+ + 2e^- \rightarrow H_2$ [1]

d) HT: Explain why this reaction is a redox reaction. (4)
Each zinc atom loses two electrons, [1] to two hydrogen ions, [1] zinc atoms oxidised as lose electrons, [1] hydrogen ions reduced as gain electrons [1]

2. a) HT: Explain why ethanoic acid (found in vinegar) is described as a weak acid, whereas nitric acid is a strong acid. (4)
Any 4 from (accept dissociate for ionise)
Ethanoic acid does not ionise completely when added to water, [1] reaction is reversible, [1] majority of molecules remain intact, [1] only a small fraction form $H^+(aq)$ ions [1]. Therefore,

ethanoic acid does not produce as high a concentration of H^+(aq) ions as a strong acid of equal concentration. [1] Nitric acid is a strong acid, because its molecules ionise completely in water [1]

b) HT: Magnesium reacts with ethanoic acid and nitric acid. What difference would you see if magnesium carbonate was reacted with ethanoic acid of the same concentration as nitric acid? (2)
Effervescence/ fizzing with ethanoic acid (1) would be slower (than with nitric acid) (1)

C. Reactions of metals part 2 – pH scale, neutralisation, salt and titration (chem)

1. Magnesium carbonate reacts with nitric acid. The equation is shown below:

$$MgCO_3(aq) + 2\ HNO_3(aq) \dashrightarrow Mg(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

Extended writing:
Plan a method to produce dry crystals of magnesium nitrate. (6)

0 marks	*Level 1 (1–2 marks)*	*Level 2 (3–4 marks)*	*Level 3 (5–6 marks)*
No relevant content	*Magnesium carbonate is added to nitric acid.*	*Excess magnesium carbonate is added to dilute nitric acid. The excess magnesium carbonate is filtered off.*	*Excess magnesium carbonate is added to dilute nitric acid. The excess magnesium carbonate is filtered off. The saturated solution is heated for a few minutes to evaporate some of the water and left in a warm place until dry crystals are formed or the crystals are dried on filter paper.*

2. i) HT: You are given a 0.50mol/dm^3 solution of nitric acid (strong) and ethanoic acid (weak). Calculate the concentration of each acid, giving your answer in g/dm^3 to 3 significant figures. (2)
31.5 g/dm^3 nitric acid (1)
30.0 g/dm^3 ethanoic acid (1)

ii) The solution of ethanoic acid has a pH of 4 and the solution of nitric acid a pH of 1. How many times greater is the concentration of H^+ ions in the nitric acid compared to the concentration in the ethanoic acid?
1,000 times greater (1)

3. **CHEMISTRY ONLY:**

a) A titration is carried out between hydrochloric acid and sodium hydroxide.
The following results show the volumes of acid added to neutralize the sodium hydroxide.

	Rough	**Trial 1**	**Trial 2**	**Trial 3**
Volume of acid added (cm^3)	15.70	15.30	15.25	15.30

Calculate the mean volume of solution added and explain your answer. (3)

(15.3 + 15.25 + 15.3) ÷ 3 = 15.28 cm^3 (2 – 1 mark for method and 1 mark for answer with unit)
Volume of rough titration has been disregarded) (1)

b) HT: In another investigation, it takes 27.00cm^3 of hydrochloric acid to neutralise 25.00cm^3 of sodium hydroxide at a concentration of 1.0 mol/dm^3. Calculate the concentration of hydrochloric acid in g/cm^3. (4)

Number of moles of sodium hydroxide = concentration × volume

= 1 mol/dm^3 × (25 ÷ 1000) dm^3 = 0.025 mol (1)

The equation for the reaction shows that 1 mole of sodium hydroxide reacts with 1 mole of hydrochloric acid. So there is 0.025 mol of HCl in 27 cm^3 of solution.

So the concentration of HCl in mol/dm^3 = number of moles ÷ volume

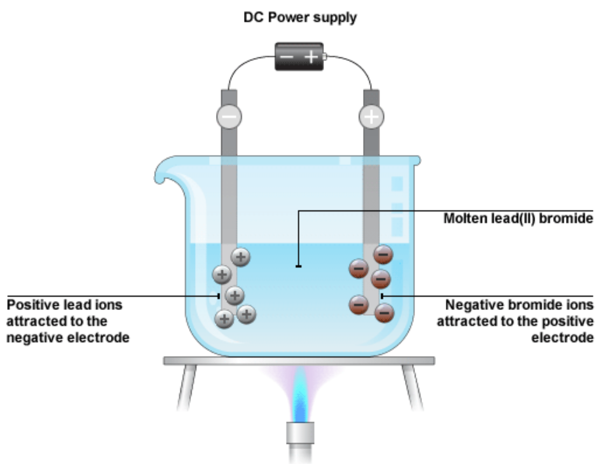
= 0.025 mol ÷ (27 ÷ 1000) dm^3 = 0.925 mol/dm^3 (1)

The mass of 1 mole of HCl is (1 + 35.5) = 36.5 g (1)

So the concentration in g/dm^3 = 36.5 g/mol × 0.925 mol/dm^3 = 33.8 g/dm^3 (1)

D. Electrolysis part 1 – Electrolysis of a molten and solution

1. The diagram shows how molten lead bromide is electrolysed.
Lead bromide contains Pb^{2+} and Br^- ions.



a) Explain why molten lead bromide conducts electricity. (1)

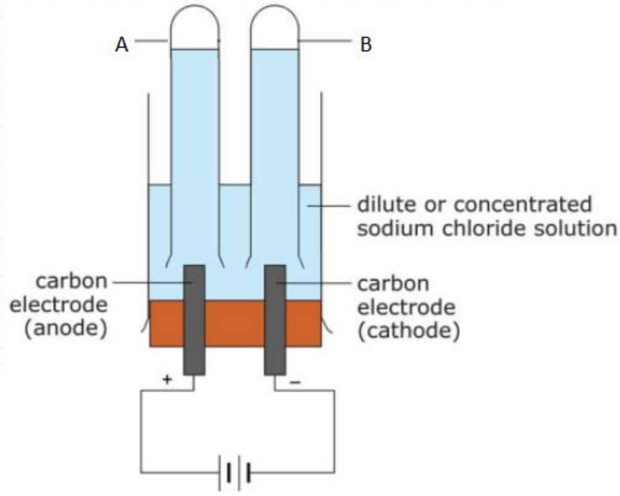
The ions are free to move/carry the electrical charge (1)

b) HT: Write the half equations, including the state symbols for the changes at the anode and cathode. (4)

$2Br^- \rightarrow Br_2 + 2e^-$ (2)

$Pb^{2+} + 2e^- \rightarrow Pb$ (2)

2. The diagram shows how sodium chloride is electrolysed in the laboratory:



a) Name the products A and B? (2)

A: chlorine (1)

B: hydrogen (1)

b) Give one use of substance A. (1)

To make bleach/plastics/used for sterlising water (any one for 1 mark)

c) A few drops of universal indicator was added to the solution after the reaction and it turned blue. Explain why. (2)

The solution is alkaline (1) because sodium hydroxide (ions) is produced (1)

d) HT: Write the half equations, including the state symbols for the changes as the anode and cathode. (4)

anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$ or $2Cl^-(aq) - 2e^- \rightarrow Cl_2(g)$ (2)

cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (2)

E. Electrolysis part 2 – Using electrolysis to extract metals

1. Aluminium is extracted from Aluminium oxide (Al_2O_3) by electrolysis. Aluminium contains Al^{3+} and O^{2-} ions.

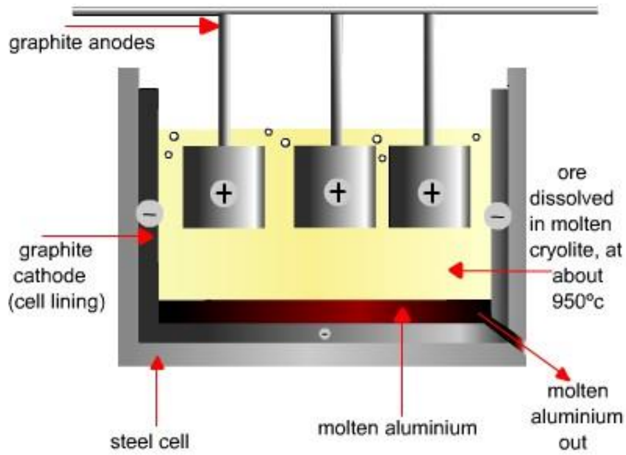
a) Suggest why aluminium was only discovered in the 1800s, despite it being a common element in the Earth's crust. (3)

Aluminium in its ore is bonded to other elements in compounds difficult to break down. [1]

Aluminium could only be extracted once electrical cells used to pass electricity were made [1]

Aluminium could then be made from one of its molten compounds by electrolysis. [1]

The following diagram shows how aluminium is extracted from aluminium oxide by electrolysis:



b) Why is molten aluminium oxide dissolved in molten cryolite? (2)

To melt at a lower temperature (1) and therefore save on costs (1)

c) Why are the carbon anodes replaced regularly in the industrial electrolysis of aluminium oxide? (2)

Any two from

Oxygen produced at the hot carbon anodes [1] reacts with the carbon to produce carbon dioxide, burning away anodes. [1]

d) HT: Write half equations for the changes at each electrode and explain which of the ions are oxidised and reduced. (4)

cathode: $Al^{3+} + 3e^{-} \rightarrow Al$ [1] Al^{3+} ions gain electrons - reduced [1]

anode: $2O^{2-} \rightarrow O_2 + 4e^{-}$ [1] O^{2-} ions lose electrons - oxidised [1]