# Explanatory Notes: GCSE Chemistry Paper 1H (May/June 2019)

These notes are designed to do more than just give you the answers; they will break down each question and explain the core chemistry principles behind them. By understanding the 'why', you can build the knowledge and confidence to tackle similar problems in the future.

## 1. Question 01: Reactions of Metals & Reactivity

This section focuses on key concepts like the reactivity series, how to design a fair experiment, and how reactivity changes based on an element's position in the periodic table.

### 1.1. Reactivity Order (Question 01.1)

- **Correct Answer:** `Ca Mg Zn Cu` (in decreasing order of reactivity).
- **The 'Why':** In Figure 1 of the exam paper, the reactivity of a metal is shown by how vigorously it reacts with acid. This is observed by the rate of fizzing, which is the production of hydrogen gas.
  - **Calcium (Ca)** and **Magnesium (Mg)** fizz the most, indicating they are the most reactive.
  - **Zinc (Zn)** fizzes less vigorously.
  - **Copper (Cu)** shows no reaction, making it the least reactive.

### 1.2. Controlling Variables (Question 01.2)

- **Key Variables:** Two variables that must be kept constant are:
  - Concentration of the acid
  - Volume of the acid
  - *(Other valid answers include the surface area of the metal or the temperature of the acid.)*
- **The 'Fair Test' Principle:** These factors are **control variables**. In any scientific investigation, you must keep control variables the same to ensure a 'fair test'. This guarantees that the only thing affecting the outcome (the rate of reaction) is the variable you are intentionally changing (the type of metal).

### 1.3. The Independent Variable (Question 01.3)

- **Answer:** The `type of metal`.
- **Definition:** The **independent variable** is the one factor that a scientist deliberately changes or manipulates during an experiment to observe its effect on the outcome.

### 1.4. Group 2 Reactivity Trend (Question 01.4)

- **Prediction:** Beryllium is **less reactive** than magnesium.
- **Scientific Reason:** Reactivity increases as you go *down* Group 2. Beryllium is above magnesium, so it is less reactive. This trend is explained by atomic structure:
    - **Atomic Size:** Beryllium atoms are smaller than magnesium atoms.
    - **Electron Shells & Shielding:** Beryllium has fewer electron shells. This means there is less 'shielding' by inner electrons to block the pull of the nucleus on the outer electrons.
    - **Electrostatic Attraction:** The combination of a smaller size and less shielding results in a stronger electrostatic force of attraction between the positive nucleus and the negative outer electrons.
    - **Conclusion:** Because of this strong attraction, more energy is needed to remove the outer electrons for a reaction to occur, making beryllium less reactive than magnesium.

### 1.5. Concentration Calculation (Question 01.5)

- **Final Answer:** `64 g/dm³`
- **Step-by-Step Calculation:**
    1. **Convert Volume:** The volume is given in cm³, but the concentration unit requires dm³. To convert, you divide by 1000. `50 cm³ / 1000 = 0.05 dm³`
    2. **Use the Formula:** The formula for concentration is: `Concentration = Mass (g) / Volume (dm³)`
    3. **Calculate:** Substitute the values into the formula. `Concentration = 3.2 g / 0.05 dm³ = 64 g/dm³`

Understanding how to predict reactivity and control experiments are fundamental skills for any chemist.

## 2. Question 02: Salts, Acids & Alkalis

This section covers the properties of acids and alkalis, the process of neutralisation, and related quantitative chemistry calculations.

### 2.1. State Symbols (Question 02.1)

- **Answer:** `(aq)`
- **Meaning:** The state symbol `(aq)` stands for **aqueous solution**. It is used for any substance that is dissolved in water. Since nitric acid is used as a solution in this reaction, `(aq)` is the correct symbol.

### 2.2. Chemical Formulae (Question 02.2)

- **Answer:** $HNO_3$ is the correct chemical formula for nitric acid.

## 2.3. Universal Indicator Colours (Question 02.3)

- **Colours:**
  - In nitric acid: `red`
  - In ammonia solution: `blue` or `purple`
- **The pH Scale Link:** Universal indicator is used to estimate the pH of a solution. Nitric acid is a strong acid with a low pH (typically 1-2), which turns the indicator red. Ammonia solution is a weak alkali with a pH above 7 (typically 10-11), which turns the indicator blue or purple.

## 2.4. Neutralisation and Excess Acid (Question 02.4)

- **Correct Row:** `D`
- **Explanation:**
  1. **Starting pH:** Ammonia solution is a weak alkali, so its pH will be above 7. A value of 10 is a reasonable starting point.
  2. **Final pH:** The question states that an *excess* of acid is added. This means that after all the ammonia has been neutralised (reaching pH 7), more acid is added. This makes the final solution acidic, so its pH must be less than 7. A value of 2 is appropriate for an excess of strong acid.

## 2.5. Percentage by Mass Calculation (Question 02.5)

- **Final Answer:** `60%`
- **Step-by-Step Calculation:**
  1. **Identify Atomic Masses (Ar):** From the question, N=14, H=1, O=16.
  2. **Formula Mass (Mr):** The Mr of ammonium nitrate ($NH_4NO_3$) is given as 80.
  3. **Find Total Mass of Oxygen:** The formula $NH_4NO_3$ contains three oxygen atoms. `Total Mass of O = 3 × 16 = 48`
  4. **Use the Formula:** The formula for percentage by mass is: `(Total Ar of element / Mr of compound) × 100`
  5. **Calculate:** `(48 / 80) × 100 = 60%`

## 2.6. Required Practical Method (Question 02.6)

- **Key Steps:** To investigate the temperature change when different masses of ammonium nitrate are dissolved in water, you should:
  1. Measure a fixed volume of water (e.g., 25 $cm^3$) and pour it into a **polystyrene cup**. A polystyrene cup is used because it is a good insulator, which minimises heat exchange with the surroundings.
  2. Record the **initial temperature** of the water using a thermometer.
  3. Add a known mass of ammonium nitrate (e.g., 2 g).
  4. **Stir** the solution gently with the thermometer to ensure the solid dissolves completely.
  5. Watch the thermometer and record the final (lowest or highest) temperature reached.

6. To investigate the effect of *different masses*, the experiment must be repeated. For each repeat, use the same volume of water but change the mass of ammonium nitrate added (e.g., 4 g, 6 g, etc.).

These questions demonstrate how chemists use both qualitative observations, like colour changes, and precise quantitative calculations to understand chemical reactions.

# 3. Question 03: Energy Changes & Bonding

This section explores the energy changes that occur during chemical reactions and the fundamental principles of how atoms bond together.

## 3.1. Reaction Profiles (Question 03.1 & 03.2)

- **Activation Energy:** The **activation energy** is the minimum amount of energy required to start a reaction. On a reaction profile, it is shown as an arrow drawn from the energy level of the reactants (`800 kJ`) up to the peak of the curve.
- **Overall Energy Change:**
    - **Calculation:** The overall energy change is `-500 kJ`.
    - **The 'Why':** This is calculated by finding the difference between the energy of the products and the energy of the reactants. `Energy Change = Energy of Products - Energy of Reactants` `Energy Change = 300 kJ - 800 kJ = -500 kJ`
    - The **negative sign** is crucial. It signifies that energy has been released to the surroundings, meaning this is an **exothermic** reaction. You can see this visually on the graph because the products are at a lower energy level than the reactants.

## 3.2. Covalent Bonding (Question 03.3)

- **Covalent Bonding in Oxygen:** An oxygen atom is in Group 6, so it has 6 electrons in its outer shell. To achieve a stable, full outer shell of 8 electrons, it needs to gain two more. It does this by **sharing** two of its electrons with another oxygen atom.
- **Dot and Cross Diagram:** The diagram for an $O_2$ molecule should show:
    - **Two shared pairs** of electrons (one dot and one cross in each pair) in the overlapping region between the two atoms. This is a **double covalent bond**.
    - **Four non-bonding electrons** on the outer shell of each individual oxygen atom.

## 3.3. Bond Energy Calculation (Question 03.4)

- **Final Answer:** `-220 kJ`
- **Step-by-Step Calculation:** The overall energy change is found using the formula: **Energy Change = Sum of bonds broken − Sum of bonds made**
    1. **Bonds Broken (Reactants):** We start with `2 H-O-O-H` molecules. This means we must break:

        - 4 × O-H bonds
        - 2 × O-O bonds
        - Energy IN = (4 × 463 kJ) + (2 × 138 kJ) = 1852 + 276 = 2128 kJ
2. **Bonds Made (Products):** We form 2 H-O-H and 1 O=O molecule. This means we make:
        - 4 × O-H bonds
        - 1 × O=O bond
        - Energy OUT = (4 × 463 kJ) + 496 kJ = 1852 + 496 = 2348 kJ
3. **Overall Change:**
        - Energy Change = 2128 kJ - 2348 kJ = -220 kJ

Energy changes in reactions are fundamentally linked to the breaking and making of chemical bonds, which are determined by an element's atomic structure.

# 4. Question 04: The Periodic Table

This section covers the historical development of the periodic table and explores the properties and trends of elements in Group 7 (the Halogens) and Group 0 (the Noble Gases).

## 4.1. History of the Periodic Table (Question 04.1 & 04.2)

- **Early Arrangement:** The earliest periodic tables arranged the elements in order of their **atomic weight**.
- **Mendeleev's Contribution:** Dmitri Mendeleev made a crucial breakthrough. He improved the table by:
    - **Leaving gaps** for elements he predicted existed but had not yet been discovered.
    - Sometimes swapping the order of elements (ignoring strict atomic weight) to ensure they fitted into groups with similar chemical properties.

## 4.2. Group 7 Properties and Trends (Question 04.3 & 04.4)

- **Low Boiling Points:** The halogens (Group 7) exist as **simple molecules** (e.g., $Cl_2$). The forces of attraction *between* these individual molecules are **weak intermolecular forces**. Because these forces are weak, they require very little energy to overcome, which results in low boiling points.
- **Trend in Boiling Points:** The boiling points of the halogens *increase* as you go down Group 7.
    - As you descend the group, the atoms and therefore the molecules get larger.
    - This increase in size leads to stronger intermolecular forces between the molecules.

   - More energy is needed to overcome these stronger forces, resulting in higher boiling points.

### 4.3. Group 0 (Noble Gases) (Question 04.5)

- **Unreactivity:** Neon is unreactive because it has a **stable electron arrangement**.
- **Electronic Structure:** Its electronic structure is **2,8**. This shows that it has a **full outer shell** of electrons. Because its outer shell is full, neon has no tendency to lose, gain, or share electrons to form chemical bonds, making it chemically inert (unreactive).

### 4.4. Moles and Avogadro's Constant (Question 04.6)

This principle of quantifying elements applies to all atoms, including other noble gases like argon.

- **Final Answer:** $1.51 \times 10^{22}$ atoms.
- **Step-by-Step Calculation:**
  1. **Calculate Moles:** First, use the mass and the relative atomic mass (Ar) of argon to find the number of moles.
     - `Moles = Mass (g) / Ar`
     - `Moles = 1 g / 40 = 0.025 mol`
  2. **Calculate Atoms:** Next, multiply the number of moles by the Avogadro constant ($6.02 \times 10^{23}$) to find the number of individual atoms.
     - `Number of Atoms = Moles × Avogadro Constant`
     - Number of Atoms = $0.025 \times (6.02 \times 10^{23}) = 1.51 \times 10^{22}$

The periodic table is an incredibly powerful tool that not only organizes elements but also helps us predict their properties and quantify matter using concepts like the mole.

## 5. Question 05: Electrolysis

This section covers the industrial process of electrolysis, its use in extracting reactive metals, and the skills needed to interpret experimental data and graphs.

### 5.1. Principles of Electrolysis (Question 05.1 & 05.2)

- **Reason for Electrolysis:** Electrolysis is used to extract metals that are **more reactive than carbon**. Carbon is not reactive enough to displace these metals from their ores, so electricity must be used to force the decomposition.
- **Electrolysis of Aluminium Oxide:** The molten mixture used to produce aluminium contains two substances:
  1. **Aluminium oxide ($Al_2O_3$):** This is the compound that contains the aluminium.
  2. **Cryolite:** Aluminium oxide has an extremely high melting point (over 2000°C). Molten cryolite is used as a solvent to dissolve the aluminium oxide, which

**lowers the operating temperature** to around 950°C. This dramatically reduces the energy required and makes the process economically viable.

## 5.2. Half Equations (Question 05.3)

- **The Process:** In the electrolysis of a molten ionic compound (like copper chloride), the positive metal ions ($Cu^{2+}$) are attracted to the negative electrode (cathode), where they gain electrons (reduction). The negative non-metal ions ($Cl^-$) are attracted to the positive electrode (anode), where they lose electrons (oxidation).
- **Half Equations:**
  - **Negative electrode (cathode):** $Cu^{2+} + 2e^- \rightarrow Cu$
  - **Positive electrode (anode):** $2Cl^- \rightarrow Cl_2 + 2e^-$

## 5.3. Interpreting Experimental Data (Question 05.4, 05.5, 05.6, 05.7)

- **Practical Errors (05.4):** The measured mass of copper might be different from the expected value for several reasons, including:
  - Some of the deposited **copper flaked off the electrode** before the final weighing.
  - The **electrode was not fully dry** when it was weighed, meaning the mass of residual water was also included.
- **Graph Extrapolation (05.5):** To find the mass produced in 24 hours from Figure 5, you can scale up a reading from the graph.
  - Read a convenient value from the line, e.g., at `10 minutes`, the mass produced is `1.8 mg`.
  - Convert 24 hours into minutes: `24 hours × 60 mins/hour = 1440 minutes`.
  - Scale up the mass: `(1440 minutes / 10 minutes) × 1.8 mg = 259.2 mg`.
- **Reading the Y-intercept (05.6):** The mass of the electrode at the start of the experiment (at time = 0) is the point where the line crosses the y-axis (the y-intercept). Reading from Figure 6, this value is `4.75 g`.
- **Calculating Gradient (05.7):**
  - **Method:** The gradient of a straight line represents the rate of change and is calculated as the change in the y-axis divided by the change in the x-axis (`change in y / change in x`).
  - **Calculation:** First, identify two points that are clearly on the line in Figure 6. The line starts at (0 hours, 4.75 g) and ends at (10 hours, 6.75 g).
    - Change in y (mass) = `6.75 g - 4.75 g = 2.0 g`
    - Change in x (time) = `10 hours - 0 hours = 10 hours`
    - Gradient = `2.0 g / 10 hours = 0.2 g/hour`
  - **Unit:** The unit is the y-axis unit divided by the x-axis unit, which is grams per hour or `g/hour`.

Electrolysis is not just a theoretical concept but a practical industrial process whose efficiency can be measured, analysed, and improved using data.

# 6. Question 06: Group 1 Metals & Bonding

This section contrasts the properties of highly reactive Group 1 metals with the different types of chemical bonds they can form with non-metals.

## 6.1. Balancing Equations & Observations (Question 06.1 & 06.2)

- **Balanced Equation:** $2Na + Cl_2 \rightarrow 2NaCl$
- **Reaction Observations:**
    - **Before:** The reaction starts with a `silvery metal` (sodium) and a `pale green gas` (chlorine).
    - **During:** When the hot sodium reacts with chlorine, a `bright yellow flame` is observed.
    - **After:** The product of the reaction is a `white solid` called sodium chloride.

## 6.2. Group 1 Reactivity Trend (Question 06.3)

- **The Trend:** Sodium is **less reactive** than potassium. The reactivity of Group 1 metals increases as you go down the group. This is because:
    - Potassium has more electron shells than sodium, meaning its single outer electron is **further from the nucleus**.
    - There is more **shielding** of the nuclear charge by the inner electron shells in potassium.
    - This combination of greater distance and increased shielding results in a **weaker force of attraction** between the positive nucleus and the negative outer electron.
    - Because the attraction is weaker, potassium's outer electron is **lost more easily**, making it more reactive than sodium.

## 6.3. Comparing Ionic and Covalent Compounds (Question 06.4)

- **The 6-Mark Question:** This question requires a detailed comparison of both the structure and the bonding in sodium chloride (an ionic compound) and hydrogen chloride (a covalent compound).

| Feature | Sodium Chloride (NaCl) | Hydrogen Chloride (HCl) |
| --- | --- | --- |
| **Type of** | **Ionic bonding**. | **Covalent bonding**. |

| Bonding | | |
|---|---|---|
| **How it Forms** | Formed by the **transfer** of an electron from a metal (Na) to a non-metal (Cl), creating oppositely charged ions ($Na^+$ and $Cl^-$). | Formed by the **sharing** of a pair of electrons between two non-metals (H and Cl). |
| **Structure** | A **giant ionic lattice**—a regular, repeating 3D arrangement of positive and negative ions. | Exists as **simple, discrete molecules**. |
| **Forces** | **Strong electrostatic forces** of attraction between the oppositely charged ions, acting in all directions. | **Weak intermolecular forces** *between* the individual HCl molecules. (The covalent bond *within* the molecule is strong). |