MEGA LECTURE

AQA (GCSE Notes)

Chapter 1: Atomic Structure and the Periodic Table

Q1. What is an atom and how is it related to an element?
Answer: An atom is the smallest part of an element that can exist while still retaining the properties of that element. Each element is made up of only one kind of atom. For example, all the atoms in a gold element are gold atoms. Atoms combine in different ways to form compounds, but the element itself is made from identical atoms.

Q2. Why is an atom considered the smallest part of an element?
Answer: An atom is considered the smallest part of an element because it cannot be broken down into anything simpler using chemical methods and still retain the identity of that element. Once you split an atom, it no longer behaves like that element. So, it is the basic building block of all substances.

Q3. What does the symbol 'O' represent in terms of atomic structure?
Answer: The symbol 'O' represents an atom of oxygen. It tells us the type of element and is a shorthand used in chemical equations and formulae. In atomic structure, an oxygen atom has 8 protons, 8 electrons, and usually 8 neutrons. The symbol helps identify the element without needing to write the full name.

Q4. What is meant by a chemical symbol and give two examples?
Answer: A chemical symbol is a one- or two-letter abbreviation that represents an element. It is used in chemical formulae and equations. For example, 'H' stands for hydrogen and 'Na' stands for sodium. These symbols make it easier to write chemical information clearly and concisely.

Q5. How many elements are approximately found in the periodic table?
Answer: There are approximately 100 different elements listed in the periodic table. Each element has a unique atomic number and chemical symbol. These elements make up all the substances in the universe through different combinations and chemical bonding.

Q6. What is the periodic table and what does it show?
Answer: The periodic table is a chart that organizes all known chemical elements based on their atomic number and properties. It shows the elements in groups and periods, indicating similar chemical behaviour in groups. It helps predict how elements will react and form compounds.

Q7. What is a compound and how is it formed?
Answer: A compound is a substance made from two or more different elements that are chemically combined in fixed proportions. Compounds are formed through chemical reactions where atoms bond together, usually by sharing or transferring electrons, to create a new substance with different properties from the original elements.

Q8. Why do chemical reactions lead to the formation of new substances?
Answer: Chemical reactions rearrange the atoms of the substances involved, breaking old bonds and forming new ones. This results in the formation of new substances with different chemical and physical properties. Energy is often absorbed or released during these changes, showing that a new substance has been made.

Q9. What kind of energy changes can be observed during chemical reactions?
Answer: Chemical reactions often involve energy changes. These can include heat, light, sound, or electrical energy. For example, in exothermic reactions, energy is released as heat, while in endothermic reactions, energy is absorbed from the surroundings. These changes can be observed through temperature rise or drop.

Q10. Why must compounds have fixed proportions of elements?
Answer: Compounds must have fixed proportions of elements because they form through specific chemical bonds where atoms combine in set ratios. This ensures the compound has consistent chemical properties. Changing the ratio would result in a different substance with different characteristics.

Q11. How are compounds represented using chemical symbols?
Answer: Compounds are represented using chemical formulae that show the symbols of the elements involved and the ratio in which their atoms combine. For example, H_2O represents water, showing two hydrogen atoms and one oxygen atom. These formulae help understand the composition of the compound easily.

Q12. Why can't compounds be separated into elements by physical processes?
Answer: Compounds can't be separated into elements by physical processes because the elements are chemically bonded. Physical methods like filtering or distilling do not break chemical bonds. Only chemical reactions can separate a compound into its original elements by breaking these bonds.

Q13. Describe how a chemical reaction can be represented using a word equation.
Answer: A word equation uses the names of the reactants and products to show what happens in a chemical reaction. It follows the pattern: Reactants → Products. For example, "Hydrogen + Oxygen → Water" shows that hydrogen and oxygen react to form water. It is a simple way to show chemical changes.

Q14. What is the difference between a word equation and a symbol equation?
Answer: A word equation uses full names of chemicals, while a symbol equation uses chemical symbols and formulae. For example, "Hydrogen + Oxygen → Water" is a word equation, and "$2H_2 + O_2 \rightarrow 2H_2O$" is a symbol equation. Symbol equations give more detail and show the number of atoms involved.

Q15. Give an example of a chemical formula for a compound and explain what it shows.
Answer: An example of a chemical formula is CO_2, which stands for carbon dioxide. It shows that

each molecule contains one atom of carbon and two atoms of oxygen. This formula also implies a fixed proportion and gives information about the elements and their quantities in the compound.

Q16. What does it mean to balance a chemical equation?
Answer: Balancing a chemical equation means making sure that the number of atoms of each element is the same on both sides of the equation. This is important because atoms are neither created nor destroyed in a chemical reaction. Balanced equations follow the law of conservation of mass.

Q17. What is the significance of using correct chemical symbols when writing formulae?
Answer: Using correct chemical symbols ensures that the formula accurately represents the substances involved. It avoids confusion and errors in understanding or predicting chemical reactions. A small mistake in a symbol can refer to a completely different element or compound, so accuracy is essential.

Q18. Write the symbol for an atom of sodium and an atom of chlorine.
Answer: The symbol for an atom of sodium is Na, and the symbol for an atom of chlorine is Cl. These symbols are used in chemical equations and formulae to represent the elements involved in a reaction or compound.

Q19. Name the compound formed when magnesium reacts with oxygen.
Answer: When magnesium reacts with oxygen, the compound formed is called magnesium oxide. Its chemical formula is MgO. This reaction involves magnesium atoms each bonding with one oxygen atom to form a white powdery compound commonly seen when magnesium burns.

Q20. Why is it important to learn the names and symbols of the first 20 elements?
Answer: Learning the names and symbols of the first 20 elements helps in understanding basic chemistry. These elements often appear in chemical reactions, formulas, and equations. Knowing them supports correct interpretation of scientific information and helps solve problems in exams and real-life situations.

Q21. What are Group 1 elements known as and give one of their properties?
Answer: Group 1 elements are known as alkali metals. One of their properties is that they are very reactive, especially with water, forming alkaline solutions and hydrogen gas. They also have low melting points compared to other metals and are soft enough to be cut with a knife.

Q22. What are Group 7 elements known as and give one of their properties?
Answer: Group 7 elements are called halogens. One of their properties is that they are reactive non-metals that form salts when combined with metals. Their reactivity decreases as you go down the group, and they exist in different physical states at room temperature, like gases or solids.

Q23. How would you name a compound formed from potassium and bromine?
Answer: A compound formed from potassium and bromine would be called potassium bromide. The

name of the metal (potassium) stays the same, and the non-metal (bromine) changes its ending to "-ide." Its chemical formula is KBr, showing it contains one potassium atom and one bromine atom.

Q24. Write a word equation for the reaction between hydrogen and oxygen to form water.
Answer: The word equation for this reaction is: Hydrogen + Oxygen → Water. This shows that hydrogen reacts with oxygen to produce water. It is a simple and clear way to describe the chemical reaction without using symbols or formulas.

Q25. Write the balanced chemical equation for the reaction between sodium and chlorine.
Answer: The balanced chemical equation is: $2Na + Cl_2 \rightarrow 2NaCl$. This shows that two atoms of sodium react with one molecule of chlorine (which has two atoms) to form two units of sodium chloride. The equation is balanced with equal numbers of each type of atom on both sides.

Q26. What is a half equation and when is it used?
Answer: A half equation shows the movement of electrons in a redox reaction. It focuses on one species either gaining or losing electrons. It is used to show what happens at each electrode in electrolysis or during displacement reactions. For example, $Na \rightarrow Na^+ + e^-$ shows sodium losing one electron.

Q27. What is an ionic equation and how does it differ from a full equation?
Answer: An ionic equation shows only the reacting ions and the products they form. It removes spectator ions which do not change. A full equation shows all substances, including those that do not take part. Ionic equations are used to focus on the actual chemical change in reactions between solutions.

Q28. Define a mixture in simple terms.
Answer: A mixture is made of two or more substances that are not chemically joined together. These can be elements, compounds, or both. Each part of the mixture keeps its own properties, and they can usually be separated easily using physical methods.

Q29. How is a mixture different from a compound?
Answer: In a mixture, substances are simply mixed and not chemically bonded, while in a compound, the elements are chemically combined in fixed ratios. Mixtures can be separated by physical means, but compounds require chemical reactions to be split.

Q30. Why do the substances in a mixture keep their own chemical properties?
Answer: The substances in a mixture are not chemically bonded, so their particles remain unchanged. Since no new bonds are formed or broken, each substance behaves as it normally would, retaining its original chemical properties.

Q31. What is filtration used for in separating mixtures?
Answer: Filtration is used to separate an insoluble solid from a liquid. The solid is trapped in the filter paper while the liquid passes through. It's commonly used in labs to separate substances like sand from water.

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Q32. Describe the process of crystallisation and what it is used for.
Answer: Crystallisation is used to separate a dissolved solid from a solution. The solution is gently heated to evaporate the solvent. When enough liquid is gone, the solid forms crystals as it becomes less soluble. It is used to get pure solids like salt from saltwater.

Q33. When is simple distillation used to separate a mixture?
Answer: Simple distillation is used to separate a liquid from a solution, especially when the liquid has a much lower boiling point than the solute. For example, it can separate water from salty water by evaporating and then condensing the water.

Q34. What is the purpose of fractional distillation?
Answer: Fractional distillation is used to separate two or more liquids with similar boiling points. It works by heating the mixture and separating the components as they boil at different temperatures. It is used for separating substances like ethanol from water.

Q35. Describe how chromatography works in separating mixtures.
Answer: In chromatography, a mixture is placed on paper and a solvent is allowed to move through it. Different components travel at different speeds, causing them to separate. It's often used to separate and identify coloured substances like dyes or inks.

Q36. Why are physical processes used to separate mixtures?
Answer: Physical processes are used because the components in mixtures are not chemically bonded. These processes rely on differences in physical properties like boiling point, solubility, or particle size. No chemical changes happen, and no new substances are formed.

Q37. Explain why no new substances are formed during the separation of mixtures.
Answer: No new substances are formed because the components are not chemically joined. Physical separation only involves changing the form or location of substances without altering their molecular structure or chemical identity.

Q38. Suggest a method to separate salt from a saltwater solution.
Answer: The best method is crystallisation or simple distillation. In crystallisation, you gently heat the saltwater to evaporate the water, leaving salt crystals behind. In simple distillation, water is boiled off and collected, leaving salt in the flask.

Q39. Which separation technique would you use to separate two miscible liquids with different boiling points?
Answer: Fractional distillation should be used. It separates miscible liquids like ethanol and water based on their boiling points using a fractionating column to improve the separation.

Q40. What technique would you use to separate the coloured components of ink?
Answer: Chromatography is the most suitable method. It separates the colours in ink based on how fast they travel with a solvent across paper. Each dye moves differently, creating a pattern that helps identify the components.

Q41. Why is chromatography useful for identifying substances?
Answer: Chromatography separates mixtures into their individual components. Each component travels a different distance, producing a distinct pattern or spot. By comparing this to known samples, you can identify unknown substances.

Q42. Give an example of when fractional distillation is used in real life.
Answer: Fractional distillation is used in the petroleum industry to separate crude oil into useful products like petrol, diesel, and kerosene. Each product has a different boiling point and is collected at different levels in a distillation tower.

Q43. What safety measures should be taken when using a Bunsen burner during separation?
Answer: Always wear safety goggles and a lab coat. Tie back long hair and keep flammable substances away. Light the Bunsen burner with a safety flame first, then adjust. Never leave it unattended, and turn it off immediately after use.

Q44. Why is it important to use clean and dry equipment when separating mixtures?
Answer: Clean and dry equipment prevents contamination, which can affect the results. Water or leftover chemicals can interfere with the separation process or create unwanted reactions, making the results unreliable or unsafe.

Q45. What is the role of the condenser in simple distillation?
Answer: The condenser cools down the hot vapour and turns it back into liquid. As the vapour passes through the cold glass tube, it condenses into a liquid that can be collected. This step is essential for recovering the separated solvent.

Q46. Describe a step-by-step method for separating a mixture of sand and salt.
Answer: First, add water to the mixture and stir. Salt dissolves, sand does not. Filter the mixture to collect sand on the filter paper. Heat the salt solution to evaporate the water, leaving salt crystals behind. This method uses filtration and crystallisation.

Q47. What would you observe when sugar is separated from a sugar solution by crystallisation?
Answer: As the water evaporates, sugar crystals begin to form and appear as white or colourless solid shapes. These crystals grow as the solution cools or continues to evaporate. The slower the cooling, the larger and purer the crystals.

Q48. How would you separate iron filings from a mixture of iron and sulfur?
Answer: Use a magnet to attract and remove the iron filings. The magnet only pulls out the iron because it is magnetic, while sulfur is not. This is a simple physical separation based on magnetic properties.

Q49. Suggest a suitable method for purifying water in a science lab.
Answer: Simple distillation is suitable. Heat the water until it boils and turns into steam. The steam

passes through a condenser where it cools and turns back into pure water. This removes dissolved substances and impurities.

Q50. Why should separation techniques be chosen based on the properties of the substances in the mixture?
Answer: Different substances have different physical properties like solubility, boiling point, or magnetism. Choosing the right separation method ensures the process works efficiently and safely. Using the wrong method might not separate the substances properly.

Q51. What did scientists believe about atoms before the discovery of the electron?
Answer: Before the discovery of the electron, scientists believed that atoms were tiny, solid spheres that could not be divided into smaller parts. They thought atoms were the smallest unit of matter and had no internal structure or separate particles. This view treated the atom as a single, indivisible object.

Q52. What was the plum pudding model of the atom?
Answer: The plum pudding model suggested that the atom was a ball of positive charge with negatively charged electrons scattered throughout it. It was named after a dessert with raisins in it, where the positive 'pudding' represented the mass and charge, and the 'plums' were the electrons embedded in it.

Q53. How did the discovery of the electron lead to the plum pudding model?
Answer: When electrons were discovered, scientists realised that atoms contained smaller, negatively charged particles. Since the atom was still thought to be neutral overall, the plum pudding model was suggested to explain how negative electrons could exist inside a positively charged body, balancing the overall charge.

Q54. What did the plum pudding model suggest about the structure of the atom?
Answer: It suggested that the atom was a sphere of positive charge with electrons embedded evenly throughout. The model assumed that the mass and positive charge were spread out evenly across the atom, with electrons scattered like plums in a pudding.

Q55. Describe how electrons were positioned in the plum pudding model.
Answer: In the plum pudding model, electrons were thought to be fixed in place within a sea of positive charge. They were spread evenly throughout the atom, with the positive charge acting as a background to balance the negative charge of the electrons.

Q56. What is the main difference between the plum pudding model and the nuclear model?
Answer: The main difference is that the plum pudding model has no nucleus and assumes the mass and charge are spread throughout the atom, while the nuclear model shows that most of the atom's mass and positive charge are concentrated in a small central nucleus with electrons orbiting around it.

Q57. What experiment led to the rejection of the plum pudding model?
Answer: The alpha particle scattering experiment, also called the Rutherford gold foil experiment, led to the rejection of the plum pudding model. It showed that alpha particles were deflected at large angles, which could not be explained by the plum pudding model.

Q58. What happened during the alpha particle scattering experiment?
Answer: In this experiment, alpha particles were fired at a thin sheet of gold foil. Most passed straight through, but some were deflected at small angles, and a few were deflected straight back. These results showed that atoms have a small, dense, positively charged centre.

Q59. What did scientists expect to happen in the alpha scattering experiment if the plum pudding model was correct?
Answer: Scientists expected the alpha particles to pass straight through the foil with only slight deflection. This is because the plum pudding model suggested that positive charge was spread out, offering no concentrated area to cause major deflections.

Q60. What unexpected result came from the alpha particle scattering experiment?
Answer: Some alpha particles were deflected at very large angles, and a small number even bounced back. This was unexpected and could not be explained by the plum pudding model, leading to a new understanding of atomic structure.

Q61. What conclusion was made about the atom's mass based on the scattering experiment?
Answer: The experiment led to the conclusion that most of the atom's mass is concentrated in a small, dense centre called the nucleus. Since most alpha particles passed through, it also showed that the atom is mostly empty space.

Q62. What conclusion was made about the charge of the nucleus after the experiment?
Answer: Scientists concluded that the nucleus is positively charged because it repelled the positively charged alpha particles. This positive charge was strong enough to cause some of the particles to be deflected at large angles.

Q63. Why did the alpha scattering experiment lead to a new model of the atom?
Answer: The results of the experiment could not be explained by the plum pudding model. The large deflections suggested a concentrated positive charge in the atom. This led to the nuclear model, which introduced the idea of a dense nucleus at the centre.

Q64. What is the main idea behind the nuclear model of the atom?
Answer: The nuclear model says that an atom has a small, dense, positively charged nucleus at its centre, where most of the mass is concentrated. Electrons orbit this nucleus, and most of the atom is empty space, allowing particles to pass through easily.

Q65. Who suggested that electrons orbit the nucleus at specific distances?
Answer: Niels Bohr suggested that electrons orbit the nucleus at specific distances. His model

added detail to the nuclear model by explaining how electrons are arranged and why they do not spiral into the nucleus.

Q66. How did Niels Bohr improve the nuclear model?
Answer: Bohr improved the nuclear model by proposing that electrons move in fixed orbits or energy levels around the nucleus rather than randomly. This explained why atoms emit or absorb energy in specific amounts and matched experimental results.

Q67. What did Bohr's model suggest about electron movement?
Answer: Bohr's model suggested that electrons move around the nucleus in fixed paths or shells. Each shell has a set energy level, and electrons can jump between these levels by absorbing or releasing energy.

Q68. Why was Bohr's model accepted by scientists?
Answer: Bohr's model was accepted because it explained observed patterns in atomic spectra and matched experimental results. His calculations supported his ideas, and it gave a better explanation of atomic behaviour than earlier models.

Q69. How did Bohr's theoretical calculations support his model?
Answer: Bohr's calculations predicted the energy levels of electrons and matched the observed emission spectra of hydrogen. These accurate predictions gave strong support to the idea that electrons occupy specific energy levels.

Q70. What name was given to the small positive particles in the nucleus?
Answer: The small positive particles in the nucleus were named protons. Each proton carries one unit of positive charge, and they account for the positive charge and most of the mass in the nucleus.

Q71. What did later experiments reveal about the charge in the nucleus?
Answer: Later experiments showed that the positive charge in the nucleus is made up of individual units called protons. Each proton has the same amount of positive charge, and the total charge depends on the number of protons present.

Q72. Why was the name "proton" chosen for the positively charged particles?
Answer: The name "proton" comes from the Greek word "protos," meaning first or primary, to reflect its fundamental role as a building block of atomic nuclei. It was given to the positively charged particles found in the nucleus.

Q73. Who provided evidence for the neutron?
Answer: James Chadwick provided the evidence for the existence of the neutron in 1932. His work showed that there was another type of particle in the nucleus with no charge but with mass similar to that of a proton.

Q74. What role did James Chadwick play in atomic theory?
Answer: James Chadwick confirmed the existence of the neutron, which helped complete the

modern understanding of atomic structure. His discovery explained the remaining mass of the atom and led to advances in nuclear physics and atomic models.

Q75. Why was the discovery of the neutron important?
Answer: The discovery of the neutron was important because it explained why atomic mass was greater than expected based on protons alone. Neutrons also helped explain nuclear stability and allowed scientists to understand atomic nuclei more accurately.

Q76. When did James Chadwick provide evidence for neutrons?
Answer: In 1932, nearly twenty years after Rutherford's nuclear atom became accepted, English physicist James Chadwick published experiments showing that beryllium bombarded with alpha particles emitted a highly penetrating, electrically neutral radiation. By measuring how this radiation knocked protons out of paraffin wax and analysing the energies involved, Chadwick concluded that the only viable explanation was a new, uncharged particle with a mass almost identical to the proton. His paper in the Proceedings of the Royal Society persuaded the wider community, completing the basic picture of nuclear matter and earning him the 1935 Nobel Prize.

Q77. What is the charge of a neutron?
Answer: A neutron has no net electric charge; it is electrically neutral. Inside the neutron, positive and negative components balance perfectly, so it produces no overall field. This neutrality lets neutrons pass close to atomic nuclei without being repelled or attracted, making them powerful probes in nuclear experiments. Although neutral, the neutron's mass is nearly the same as a proton's, so it significantly adds to an atom's mass number. By contributing mass without charge, neutrons stabilise the nucleus: their presence reduces repulsion between protons and helps determine which isotopes are stable, radioactive, or prone to fission.

Q78. Why did it take 20 years after the nucleus was discovered to find the neutron?
Answer: Detecting an uncharged particle was far harder than spotting charged ones, because early detectors relied on electric or magnetic deflection and ionisation tracks. Neutral radiation leaves almost no direct trace, so signals from neutrons were easily mistaken for high-energy gamma rays. Instruments such as cloud chambers, Geiger counters, and vacuum pumps had to improve before subtle recoil effects could be resolved. Theoretical hints also mattered: scientists first had to suspect a missing neutral mass to look for it. Only after stronger radioactive sources, better shielding, and precise energy measurements became available could Chadwick show that mysterious neutral radiation possessed mass and momentum consistent with a new particle, the neutron.

Q79. What does the modern atomic model include that earlier models did not?
Answer: The modern quantum mechanical model keeps the tiny, dense nucleus but adds protons, neutrons, and electrons arranged in probabilistic orbitals rather than fixed paths. It explains electron energy levels with quantum numbers, incorporates subshells (s, p, d, f), spin, and the Pauli exclusion principle, and predicts chemical behaviour from electron configuration. It also accounts for isotopes by varying neutron numbers and describes nuclear forces, radioactivity, and the wave–particle duality

of electrons—concepts absent from solid-sphere, plum pudding, or simple orbital models. Overall, it links atomic structure, spectra, bonding, and periodic trends in one coherent framework.

Q80. Why do scientific models change over time?
Answer: Scientific models change because new evidence, often produced by better instruments or novel experiments, reveals shortcomings in older explanations. When observations contradict predictions, scientists revise or replace the model, keeping parts that still fit. Peer review, replication, and open debate ensure that only robust ideas survive. As technology advances—from particle accelerators to space telescopes—researchers collect more precise data, forcing adjustments. This self-correcting cycle, central to the scientific method, transforms simple early models into richer, more accurate descriptions of nature while remaining open to future refinement.

Q81. What is meant by a scientific model?
Answer: A scientific model is a simplified representation—verbal, visual, or mathematical—that captures the essential features of a system so people can understand, predict, and communicate its behaviour. Like a map, it omits details that are not needed for its purpose. Models range from physical globes to equations describing gases. They generate testable predictions, guide new experiments, and are judged by how well they match evidence. Because they are tools rather than final truths, models are continually refined or replaced when better explanations emerge.

Q82. Give an example of how experimental evidence can change scientific ideas.
Answer: The shift from the geocentric to the heliocentric view of the solar system is a classic example. Precise observations by Tycho Brahe showed that planetary motions did not fit Earth-centred predictions. Johannes Kepler's analysis revealed that planets travel in ellipses around the Sun, and Galileo's telescope revealed moons orbiting Jupiter, proving that not everything circles Earth. These pieces of evidence undermined the long-held Ptolemaic system and persuaded scientists to adopt the Sun-centred model, illustrating how strong experimental data can overturn entrenched beliefs.

Q83. Why is it important to test scientific theories with experiments?
Answer: Experiments provide objective checks on whether a theory's predictions match reality. Without testing, ideas could drift into unsupported speculation. Controlled experiments reveal hidden variables, quantify relationships, and expose errors or biases. Reproducible results allow other researchers to confirm findings, building confidence in conclusions. By challenging theories, experiments drive refinement, ensuring that accepted explanations remain grounded in observable facts and yielding reliable knowledge that can be applied in technology, medicine, and everyday life.

Q84. What role does experimental evidence play in science?
Answer: Experimental evidence is the cornerstone of science. It supplies the measurable facts that theories must explain and sets quantitative limits on acceptable ideas. Evidence can confirm hypotheses, rule out alternatives, or highlight anomalies that spark new research. Because good evidence is observable, repeatable, and peer-reviewed, it provides a common standard that scientists

worldwide can trust. By anchoring theories to reality, evidence guards against bias and ensures that knowledge grows on a solid empirical foundation.

Q85. Why is the development of atomic models a good example of how science works?
Answer: The atomic model story shows science's iterative nature: each model—solid sphere, plum pudding, nuclear, Bohr, quantum—was based on the best evidence available at the time. New experiments such as electron discovery, alpha scattering, and spectroscopy revealed flaws and inspired revised models that better fit the data. Debate, peer review, and replication ensured rigorous scrutiny. This evolution demonstrates key principles of science: reliance on evidence, willingness to change ideas, and cumulative improvement of understanding over time.

Q86. What was the main limitation of the plum pudding model?
Answer: The plum pudding model could not account for the sharp deflections and occasional backward rebounds of alpha particles in Rutherford's gold-foil experiment. If positive charge were spread evenly through the atom, heavy, fast alpha particles should have sailed through with minimal disturbance. The observed large-angle scattering indicated a concentrated positive core, contradicting the model. It also offered no explanation for discrete spectral lines, nor did it specify how electrons were held in place, making it inadequate once precise measurements were available.

Q87. How did the nuclear model explain the deflection of alpha particles?
Answer: The nuclear model proposed that nearly all an atom's mass and positive charge are packed into a minute central nucleus. Alpha particles passing close to this dense, positively charged centre experience strong electrostatic repulsion, causing sharp bends or even rebounds. Because the nucleus is tiny compared with the whole atom, only a small fraction of particles approach it closely, so most pass straight through. This arrangement neatly explained both the rare, large deflections and the common, minor ones observed in Rutherford's experiment.

Q88. How did the discovery of subatomic particles change the idea of an atom?
Answer: Finding electrons, protons, and neutrons showed that atoms are not indivisible but contain smaller components with specific roles: electrons form an outer cloud responsible for bonding; protons set the element's identity; and neutrons add mass and influence stability. This internal structure explained chemical behaviour, isotopes, radioactivity, and electrical properties. Matter was revealed to be built from interacting particles, linking chemistry and physics under shared fundamental principles rather than treating atoms as featureless solid blocks.

Q89. What are the three main subatomic particles?
Answer: The three main subatomic particles are protons, neutrons, and electrons. Protons, each with one positive charge, define the atomic number and reside in the nucleus. Neutrons carry no charge but have nearly the same mass as protons; they share the nucleus and affect nuclear stability, creating isotopes when their number varies. Electrons carry one negative charge and occupy energy levels outside the nucleus, taking part in chemical bonding. Though far lighter than nucleons, electrons fill most of the atom's volume and govern its reactivity and spectral behaviour.

Q90. Where are the protons and neutrons located in the atom?
Answer: Protons and neutrons are tightly packed together in the atom's central nucleus, a region only about 1×10^{-14} m across but containing almost all the atom's mass. Held by the strong nuclear force, these nucleons form a dense core that resists the electrostatic repulsion between protons. Because the nucleus is so small, it occupies only a tiny fraction of the atom's volume, yet its positive charge and mass dominate atomic properties and underpin processes such as radioactivity, fission, and fusion.

Q91. Where are the electrons found in the modern atomic model?
Answer: Electrons inhabit regions of space called orbitals organised into shells or energy levels around the nucleus. Instead of fixed circular paths, each orbital is a probability cloud describing where an electron is likely to be at any time. Shells fill from lower to higher energy following the Aufbau principle, with subshells labelled s, p, d, and f, and no two electrons in the same atom share the exact set of quantum numbers. This arrangement explains periodic trends, bonding capacity, magnetism, and the spectral lines atoms emit or absorb.

Q92. What is the overall charge of a neutral atom?
Answer: The overall charge of a neutral atom is zero because the number of positively charged protons in the nucleus equals the number of negatively charged electrons in its outer regions. Their charges cancel exactly. Neutrons, being uncharged, add mass but no charge. If an atom gains or loses electrons it becomes an ion, but in its natural neutral state the balance of charges allows bulk matter to exist without large static attractions or repulsions and keeps everyday objects electrically neutral.

Q93. Why did the early model describe atoms as solid spheres?
Answer: Early scientists such as John Dalton had no evidence of internal structure and found that elements combined in fixed ratios, implying indivisible units. Experimental tools of the time could not probe inside atoms, and concepts like electric charge or radiation were poorly understood. The solid-sphere idea provided a simple explanation for chemical laws like conservation of mass and definite proportions. Only later, with the discovery of electrons and advances in detecting subatomic phenomena, did evidence force a departure from the solid-sphere picture.

Q94. What was learned about atomic structure from the scattering experiment?
Answer: Rutherford's alpha scattering experiment revealed that atoms consist mostly of empty space, with a tiny, dense, positively charged nucleus at the centre. The pattern of deflections allowed scientists to estimate the nucleus's size (about 1×10^{-14} m) and confirm that its charge equals the atomic number. The experiment showed that electrons occupy the vast space around the nucleus and that the nucleus contains nearly all the atom's mass, fundamentally reshaping understanding of matter at its smallest scale.

Q95. What is the significance of most alpha particles passing straight through the gold foil?
Answer: The fact that most alpha particles passed straight through indicated that the gold atoms were mostly empty space and that their positive charge was concentrated in a very small region

rather than spread out. This observation contradicted the plum pudding model and supported the nuclear model, demonstrating that matter is not a continuous solid but largely void, with electrons occupying the volume and a dense nucleus occupying only a minute fraction.

Q96. Why were only a few alpha particles deflected?

Answer: Only alpha particles that came very close to the tiny nucleus experienced strong electrostatic repulsion and were sharply deflected; the vast majority traveled through empty space and felt little force. Because the nucleus is so small relative to the atom, the probability of a close encounter is low, explaining why large-angle deflections were rare while straight-through paths were common, matching Rutherford's calculations for a point-like positive core.

Q97. How does the modern atomic model describe electron arrangement?

Answer: Electrons are arranged in quantised energy levels called shells, each containing subshells and orbitals where electrons are likely to be found. The order of filling follows the Aufbau principle, and rules such as Hund's rule and the Pauli exclusion principle govern how electrons occupy orbitals. Electrons can move between levels by absorbing or emitting specific quanta of energy, producing characteristic spectra. This organised structure explains chemical reactivity, bonding, and periodic trends across the elements.

Q98. Why are protons and neutrons in the nucleus?

Answer: Protons and neutrons are held together in the nucleus by the strong nuclear force, which is far stronger than the repulsive electrostatic force between protons but acts only over very short ranges. This intense attraction binds nucleons into a compact cluster, providing the mass and stability necessary for atoms to exist. Electrons, which do not feel the strong force, remain outside, bound by electrostatic attraction to the positive nucleus. The balance of these forces underpins nuclear stability and radioactivity.

Q99. How does the size of the nucleus compare to the size of the whole atom?

Answer: The nucleus is about 1×10^{-14} m in radius, while the atom is roughly 1×10^{-10} m, making the atom about ten thousand times larger in diameter. In volume terms this is a difference of around one trillion: if an atom were enlarged to a stadium, the nucleus would be a pea at the centre. Despite occupying such a tiny fraction of space, the nucleus contains almost all the atom's mass, illustrating how matter is largely empty space at the atomic scale.

Q100. What charge does the nucleus have and why?

Answer: The nucleus has an overall positive charge because it contains protons, each with a single positive unit of charge, while neutrons contribute mass but no charge. The total positive charge equals the atomic number, determining the element's identity. This positive charge attracts negatively charged electrons, binding them into orbitals and allowing atoms to be electrically neutral when proton and electron numbers match. The magnitude of the nuclear charge influences ionisation energies, bonding tendencies, and the deflection of charged particles such as the alpha particles in Rutherford's experiment.

Q101. What does the term "relative mass" mean in atomic structure?
Answer: Relative mass is a way of comparing the mass of a particle, like a proton, neutron, or electron, to the mass of a proton which is taken as 1. It helps simplify calculations by avoiding very small decimal numbers. For example, protons and neutrons have a relative mass of 1, while electrons have a relative mass of around 1/1836.

Q102. Why do scientists use relative mass instead of actual mass for subatomic particles?
Answer: Scientists use relative mass because the actual mass of subatomic particles is extremely small and hard to work with. Using relative mass makes it easier to compare particles and perform calculations without using very small decimal numbers in standard units like kilograms.

Q103. How does the atomic number help in identifying an element?
Answer: The atomic number tells us the number of protons in an atom, which is unique for each element. No two elements have the same number of protons, so the atomic number is used to identify which element the atom belongs to.

Q104. Can two different elements have the same atomic number? Explain.
Answer: No, each element has a unique atomic number. If two atoms have the same atomic number, they are the same element because atomic number is defined by the number of protons in the nucleus.

Q105. What is the significance of the mass number in an atom?
Answer: The mass number shows the total number of protons and neutrons in an atom. It helps to identify isotopes and gives an idea of the atom's mass since electrons contribute very little to the overall mass.

Q106. How is an isotope represented using standard notation?
Answer: An isotope is shown with its element symbol and numbers: the mass number as a superscript and the atomic number as a subscript. For example, carbon-14 is written as $^{14}_{6}C$.

Q107. Why are the physical properties of isotopes different?
Answer: Isotopes have different numbers of neutrons, which affects their mass and physical behaviour like density and rate of diffusion. However, their chemical properties remain mostly the same because they have the same number of electrons.

Q108. Why is the electron's mass often considered negligible in calculations?
Answer: The mass of an electron is about 1/1836 of the mass of a proton or neutron. Since it's so small, it has little effect on the total mass of the atom, so it's often ignored in calculations to simplify things.

Q109. If an atom has an atomic number of 15, how many electrons does it have?
Answer: 15 electrons, because in a neutral atom, the number of electrons is equal to the number of protons, which is the atomic number.

Q110. If an atom has 20 protons and 20 neutrons, what is its mass number?
Answer: 40
Solution: Mass number = protons + neutrons = 20 + 20 = 40

Q111. What is the role of neutrons in the nucleus?
Answer: Neutrons help to hold the nucleus together by reducing repulsion between protons, which all have positive charges. They also contribute to the mass of the atom.

Q112. Why do isotopes of the same element have different mass numbers?
Answer: Isotopes have the same number of protons but different numbers of neutrons. Since mass number is the total of protons and neutrons, different neutron counts result in different mass numbers.

Q113. What unit is used to measure the size of atoms?
Answer: Nanometres (nm) are used to measure atomic size. One nanometre equals 1×10^{-9} metres.

Q114. Convert 0.1 nanometres to metres using standard form.
Answer: 1×10^{-10} metres
Solution: 0.1 nm $= 0.1 \times 10^{-9} = 1 \times 10^{-10}$ m

Q115. Why is it useful to express atomic size in standard form?
Answer: Atomic sizes are extremely small, and standard form makes it easier to write, read, and compare these tiny numbers clearly without using long strings of zeros.

Q116. What does 1×10^{-10} metres represent in terms of atomic structure?
Answer: It represents the approximate radius of an atom. Atoms are very small, and this is a typical measurement of their size.

Q117. What is the relationship between atomic radius and the size of the nucleus?
Answer: The nucleus is much smaller than the whole atom—about 1/10,000 the radius of the atom. Almost all of the atom's mass is in this tiny nucleus.

Q118. How do you determine the number of subatomic particles in an atom?
Answer: Use the atomic number to find protons and electrons (in a neutral atom), and subtract the atomic number from the mass number to find neutrons.

Q119. If an atom has a charge of 2+, what does that tell you about its electrons?
Answer: It has lost 2 electrons compared to a neutral atom. The positive charge means there are more protons than electrons.

Q120. Why do ions form from atoms?
Answer: Ions form so atoms can gain a full outer electron shell, making them more stable. This usually involves losing or gaining electrons during chemical reactions.

Q121. What happens to the number of electrons when an atom forms a negative ion?
Answer: The atom gains one or more electrons, resulting in more electrons than protons, which gives it a negative charge.

Q122. Why does the nucleus remain unchanged when an atom becomes an ion?
Answer: Ions are formed by gaining or losing electrons, which are outside the nucleus. The number of protons and neutrons in the nucleus stays the same.

Q123. What keeps the electrons in orbit around the nucleus?
Answer: The attraction between the negatively charged electrons and the positively charged nucleus (due to protons) keeps the electrons in orbit.

Q124. Why is most of the atom considered empty space?
Answer: Because the electrons are far from the nucleus relative to their size and the nucleus is extremely small. Most of the volume of the atom contains no matter.

Q125. What causes atoms to have different chemical properties?
Answer: Chemical properties are determined by the number and arrangement of electrons, especially those in the outer shell. Different elements have different numbers of electrons and electron arrangements.

Q126. How can atoms of the same element differ in mass?
Answer: Atoms of the same element can have different numbers of neutrons, which changes their mass. These are called isotopes. Since neutrons add to the mass but not to the element’s identity, isotopes of an element have the same atomic number but different mass numbers, making their masses different while keeping the same chemical behaviour.

Q127. What is the typical diameter of a nucleus in standard form?
Answer: The diameter of a nucleus is about 1×10^{-14} metres. This is much smaller than the diameter of the whole atom, which is roughly 1×10^{-10} metres. This shows that the nucleus takes up very little space in the atom, but it contains nearly all of the atom’s mass.

Q128. How much of an atom’s mass is due to electrons?
Answer: Electrons contribute almost none of the atom’s mass. Nearly all of the mass comes from protons and neutrons in the nucleus. Since an electron's mass is about 1/1836 of a proton, it is usually ignored when calculating the total atomic mass.

Q129. How many electrons are in a fluoride ion?
Answer: A fluoride ion has 10 electrons. A neutral fluorine atom has 9 electrons, equal to its number of protons. When it gains one electron to become a fluoride ion (F^-), it has 9 + 1 = 10 electrons.

Q130. If an atom of aluminium has 13 protons and 10 electrons, what is its charge?
Answer: The charge is 3+ because it has lost 3 electrons. The number of protons (13) is more than the number of electrons (10), resulting in a charge of +3.

Q131. What is the name of the particle with no charge found in the nucleus?
Answer: The particle with no charge found in the nucleus is called a neutron. Neutrons have a relative mass of 1 and help to stabilise the nucleus by reducing repulsion between the positively charged protons.

Q132. How does the number of electrons change in a positive ion?
Answer: In a positive ion, the atom loses one or more electrons. This creates more protons than electrons, resulting in an overall positive charge. The number of protons stays the same, but the electron count decreases.

Q133. What holds protons and neutrons together in the nucleus?
Answer: The strong nuclear force holds protons and neutrons together in the nucleus. This force is much stronger than the repulsive force between positively charged protons, but it only acts over very short distances inside the nucleus.

Q134. Why is it useful to use atomic models in science?
Answer: Atomic models help scientists visualise the structure of atoms, explain observations, predict chemical behaviour, and understand how particles interact. Since atoms are too small to see directly, models provide a practical way to describe and communicate ideas about atomic structure.

Q135. What information is needed to draw a nuclear symbol?
Answer: You need the atomic number (number of protons), the mass number (protons + neutrons), and the chemical symbol of the element. The mass number goes at the top left, the atomic number at the bottom left of the element symbol.

Q136. How do we know atoms are mostly empty space?
Answer: The alpha scattering experiment showed that most alpha particles passed straight through the gold foil, indicating that atoms are mostly empty space. Only a few were deflected, which showed that the positive mass is concentrated in a small nucleus.

Q137. Why do atoms of noble gases not easily form ions?
Answer: Noble gases already have full outer electron shells, so they are very stable and unreactive. They do not need to gain or lose electrons to achieve stability, which is why they rarely form ions or take part in chemical reactions.

Q138. If two atoms have the same number of protons but different electrons, are they the same element?
Answer: Yes, they are the same element because the number of protons defines the element. However, if they have different numbers of electrons, they are ions of that element. The chemical behaviour may change slightly, but the element remains the same.

Q139. What is the difference between a hydrogen atom and a hydrogen ion?
Answer: A hydrogen atom has one proton and one electron. A hydrogen ion (H^+) has only a proton

because it has lost its one electron. This makes the ion positively charged and gives it different chemical behaviour from the neutral atom.

Q140. How do you find the total number of particles in the nucleus?
Answer: You add the number of protons and neutrons. This total is called the mass number.
Solution: Mass number = atomic number (protons) + number of neutrons

Q141. Which subatomic particle determines the chemical identity of an atom?
Answer: The proton determines the chemical identity of an atom. The number of protons (atomic number) defines which element the atom is. Changing the number of protons changes the element entirely.

Q142. What does it mean if an element has a mass number of 1?
Answer: It means the total number of protons and neutrons in the atom is 1. Since hydrogen has 1 proton and usually no neutrons, hydrogen-1 is the most common isotope and has a mass number of 1.

Q143. How are atoms arranged in the periodic table?
Answer: Atoms are arranged in order of increasing atomic number. Elements with similar chemical properties are placed in the same column, called a group. Rows are called periods and show patterns in electron arrangement.

Q144. What does a nuclear model diagram show?
Answer: It shows a small, dense nucleus at the centre containing protons and neutrons, with electrons orbiting around it. This model helps explain the structure and behaviour of atoms based on experimental evidence.

Q145. How does the nuclear model explain atomic mass?
Answer: The nuclear model explains that almost all the atom's mass comes from the nucleus, which contains protons and neutrons. Electrons have very little mass, so they don't significantly affect the atomic mass.

Q146. What is the value of nano in standard form?
Answer: 1 nano (n) = 1×10^{-9}. This unit is often used to measure extremely small lengths like the size of atoms or molecules.

Q147. Why can't neutrons be used to identify elements?
Answer: Neutrons don't affect the chemical identity of an element. Different elements can have similar numbers of neutrons, and isotopes of the same element have different neutron numbers. Only the number of protons (atomic number) defines the element.

Q148. How do scientists measure the mass of subatomic particles?
Answer: Scientists use relative mass and advanced equipment like mass spectrometers to measure

the mass of subatomic particles. These tools help compare particle masses and calculate atomic masses accurately.

Q149. What is meant by the term "mass number minus atomic number"?
Answer: This gives the number of neutrons in an atom. The mass number is the total of protons and neutrons, so subtracting the atomic number (number of protons) leaves the number of neutrons.

Q150. Why do models of the atom keep improving over time?
Answer: As new experiments are done and better evidence is found, scientists update models to match what we observe. Improvements in technology and understanding lead to more accurate representations of atomic structure. This is part of how science develops over time.

Q151. What is meant by the term relative atomic mass?
Relative atomic mass is the weighted average mass of the atoms of an element compared with one-twelfth of the mass of a carbon-12 atom. It takes into account all naturally occurring isotopes and their relative abundances, allowing for a practical comparison between different elements.

Q152. Why is the relative atomic mass of an element often not a whole number?
Because it is a weighted average of the masses of all isotopes of that element, considering their natural abundances. Since isotopes have different masses and are not always present in equal amounts, the resulting average is often a decimal, not a whole number.

Q153. How do you calculate the relative atomic mass from isotope abundances?
Multiply the mass of each isotope by its percentage abundance, add all the results together, then divide by 100. This gives the weighted average mass of the element's atoms. Formula: Relative atomic mass = ($mass_1 \times abundance_1 + mass_2 \times abundance_2 + ...$) ÷ 100.

Q154. What is percentage abundance in relation to isotopes?
Percentage abundance is the proportion of a specific isotope present in a natural sample of an element. It is expressed as a percentage of the total number of atoms, and helps determine the element's relative atomic mass when combined with the isotope's mass.

Q155. Why do isotopes affect the average atomic mass of an element?
Isotopes affect average atomic mass because they have different masses due to varying numbers of neutrons. The more abundant an isotope is, the more it influences the weighted average, shifting the overall relative atomic mass closer to that isotope's mass.

Q156. What two pieces of data are needed to calculate relative atomic mass?
You need the mass number of each isotope and the percentage abundance of each isotope in a natural sample of the element. These values are used in a weighted average formula to calculate the element's relative atomic mass.

Q157. What would happen to the relative atomic mass if one isotope was more abundant?
If one isotope is significantly more abundant than others, the relative atomic mass will be closer to

the mass of that isotope. The more dominant the isotope, the more it influences the average mass of all atoms of that element.

Q158. How can you calculate the relative atomic mass when an element has two isotopes?
Multiply the mass of each isotope by its percentage abundance, add the results together, and divide by 100. This gives the weighted average. Formula: Relative atomic mass = ($mass_1 \times \%abundance_1 + mass_2 \times \%abundance_2$) ÷ 100.

Q159. If an element has three isotopes, how would you calculate the average atomic mass?
Multiply the mass of each isotope by its percentage abundance, add all the products together, and divide by 100. This gives the weighted average atomic mass, accounting for the mass and abundance of each of the three isotopes.

Q160. Why do chemists use relative atomic mass instead of actual atomic mass?
Actual atomic masses are extremely small and impractical to use in calculations. Relative atomic mass offers a more convenient, standardised way to compare elements, using carbon-12 as the reference point with a mass of exactly 12.

Q161. What is meant by electronic structure?
Electronic structure refers to how electrons are arranged in an atom's energy levels or shells. It shows how many electrons occupy each shell and helps determine an element's chemical properties and how it bonds with other atoms.

Q162. What does the electronic structure 2,8,1 tell you about an atom?
It means the atom has 2 electrons in the first shell, 8 in the second shell, and 1 in the third shell. This arrangement is typical of a sodium atom, which has 11 electrons in total, matching its atomic number.

Q163. How are electrons arranged in energy levels?
Electrons fill energy levels starting from the one closest to the nucleus. Each shell must be filled to its maximum before electrons occupy the next one. This minimises the atom's energy and results in a stable configuration.

Q164. What is the maximum number of electrons in the first energy level?
The first energy level, which is closest to the nucleus, can hold a maximum of 2 electrons. Once full, additional electrons must move to the second shell or higher energy levels.

Q165. How many electrons can be held in the second energy level?
The second energy level can hold up to 8 electrons. After the first shell is full, electrons start filling the second shell, and it must be filled before electrons enter the third level.

Q166. Why do electrons fill the lowest energy levels first?
Electrons fill the lowest energy levels first because this requires the least amount of energy. Atoms are more stable when their electrons occupy lower energy states before moving to higher ones.

Q167. How does the electronic structure relate to the period number in the periodic table?
The number of occupied energy levels (or shells) in an atom is equal to the period number in the periodic table. For example, atoms in period 3 have three occupied electron shells.

Q168. What do all elements in the same group of the periodic table have in common?
All elements in the same group have the same number of electrons in their outer shell. This gives them similar chemical properties and reactivity patterns, including how they bond and form ions.

Q169. What is the electronic structure of a magnesium atom?
Magnesium has an atomic number of 12, meaning it has 12 electrons. Its electronic structure is 2 electrons in the first shell, 8 in the second, and 2 in the third—written as 2,8,2.

Q170. What is the electronic structure of a nitrogen atom?
Nitrogen has 7 electrons in total. These are arranged as 2 electrons in the first shell and 5 in the second shell, giving it the electronic configuration 2,5.

Q171. What do the numbers in the electronic structure represent?
Each number represents the number of electrons in a particular energy level or shell, starting from the shell closest to the nucleus and moving outward. Together, they show how the electrons are distributed in the atom.

Q172. How is the number of outer electrons related to the group number?
For elements in groups 1–7, the group number equals the number of electrons in the outer shell. Group 0 elements have full outer shells, usually 8 electrons, except helium, which has only 2.

Q173. How can the electronic structure be shown in a diagram?
It's shown using concentric circles around a central nucleus. Each circle represents an energy level, and electrons are marked as dots or crosses, with each shell containing a set number of electrons.

Q174. Why are electronic structures important in predicting chemical behaviour?
They show the number of electrons in the outer shell, which determines how atoms interact in chemical reactions. Atoms with similar outer shell electrons behave similarly and tend to form specific types of bonds or ions.

Q175. What is the link between reactivity and electronic configuration?
Reactivity depends on how easily an atom can lose, gain, or share electrons to achieve a full outer shell. Atoms with nearly full or nearly empty outer shells are usually the most reactive, such as Group 1 and Group 7 elements.

Q176. How does the electronic structure of noble gases explain their lack of reactivity?
Noble gases have full outer electron shells, making them stable and unreactive. Since they already have a complete set of electrons, they do not need to gain, lose, or share electrons, which is why they rarely form bonds or participate in chemical reactions.

Q177. What happens to the electronic structure when an atom becomes an ion?
When an atom becomes an ion, it gains or loses electrons to achieve a full outer shell. Metals lose electrons to form positive ions, while non-metals gain electrons to form negative ions. This changes the number of electrons but not the number of protons.

Q178. Why do metals lose electrons and non-metals gain electrons?
Metals have few electrons in their outer shell, so it's easier to lose them and achieve a full shell. Non-metals have more outer electrons and are closer to a full shell, so they gain electrons to become more stable.

Q179. How does the periodic table arrange elements?
The periodic table arranges elements in order of increasing atomic number. Elements are placed in rows (periods) and columns (groups) based on repeating patterns in their chemical and physical properties.

Q180. Why is the periodic table called "periodic"?
It’s called "periodic" because similar chemical properties repeat at regular intervals, or periods, when elements are arranged by increasing atomic number. This repeating pattern gives the table its structure and predictive power.

Q181. What does the group number tell you about an element?
The group number indicates how many electrons are in the outer shell of an atom. For example, Group 1 elements have 1 outer electron, while Group 7 elements have 7. This number helps predict reactivity and bonding behaviour.

Q182. What does the period number tell you about an element?
The period number tells you how many energy levels or electron shells an atom has. For example, elements in period 3 have electrons in three shells. It reflects the size and energy level of the atom’s outermost electrons.

Q183. How can the position of an element in the periodic table help predict its reactivity?
An element's group shows how many outer electrons it has, which affects how easily it reacts. Its period shows how many shells it has, which influences attraction between nucleus and electrons. Together, they help predict trends in reactivity.

Q184. Which side of the periodic table contains metals?
Metals are found on the left and in the middle of the periodic table. This includes Groups 1 and 2, and the transition metals in the centre. Non-metals are on the right side, starting from Group 4 or 5 across to Group 0.

Q185. Why are elements in Group 1 so reactive?
Group 1 elements have only one electron in their outer shell, which they lose easily to become stable. The further down the group, the more easily this electron is lost due to the increasing distance from the nucleus, making them more reactive.

Q186. Why do Group 7 elements become less reactive down the group?
Group 7 elements react by gaining an electron. As you go down the group, the atoms become larger and the outer shell is farther from the nucleus, making it harder to attract an extra electron. This reduces reactivity.

Q187. What type of ions do Group 1 elements form?
Group 1 elements lose one electron to form positive ions with a +1 charge. These are called cations. For example, sodium forms Na^+ and potassium forms K^+ when they lose their single outer electron.

Q188. What type of ions do Group 7 elements form?
Group 7 elements gain one electron to form negative ions with a -1 charge. These are called anions. For example, chlorine forms Cl^- and fluorine forms F^- by gaining an extra electron to complete their outer shell.

Q189. How is the modern periodic table different from early versions?
The modern table is arranged by atomic number, not atomic mass like in earlier versions. It also leaves no gaps for undiscovered elements and groups elements more accurately based on electronic structure and chemical properties.

Q190. Why are transition metals placed in a separate block in the periodic table?
Transition metals have unique properties, such as forming coloured compounds and having variable oxidation states. They are placed in the central block (Groups 3–12) because their outer electrons occupy d-orbitals.

Q191. What property is shared by elements in Group 0?
All Group 0 elements (noble gases) have full outer electron shells, making them very stable and unreactive. They are all gases at room temperature and exist as single atoms rather than molecules.

Q192. How can you predict the properties of an unknown element using the periodic table?
By looking at the group and period of the element, you can predict its electronic structure, reactivity, state, and bonding behaviour. Elements in the same group have similar chemical properties.

Q193. What happens to the number of protons as you move across a period?
As you move across a period from left to right, the number of protons increases by one for each element. This increases the atomic number and results in a greater positive charge in the nucleus.

Q194. What happens to the number of electron shells as you go down a group?
As you go down a group, each element has one more electron shell than the one above it. This means the atoms get larger and the outer electrons are farther from the nucleus.

Q195. Why do elements in the same group have similar reactions?
They have the same number of electrons in their outer shell, which determines how they bond and react with other elements. This is why elements in the same group often form similar compounds.

Q196. What element is in Group 2 and Period 3 of the periodic table?
The element in Group 2 and Period 3 is magnesium (Mg). It has two electrons in its outer shell and three energy levels, giving it the electronic structure 2,8,2.

Q197. What element is in Group 6 and Period 2 of the periodic table?
The element in Group 6 and Period 2 is oxygen (O). It has six outer electrons and two energy levels, giving it the electronic structure 2,6.

Q198. How can you use atomic number to place an element in the periodic table?
The atomic number tells you the number of protons and electrons. Use the number of shells (for period) and outer electrons (for group) to place it. The position reflects its properties and reactivity.

Q199. Why do elements in Group 1 react violently with water?
Group 1 metals react with water to form a metal hydroxide and hydrogen gas. Their single outer electron is lost easily, and the reaction becomes more vigorous down the group as atoms get larger and the electron is more easily removed.

Q200. How can electronic structure help explain trends in reactivity?
Electronic structure shows how many outer electrons an atom has and how easily they can be lost or gained. Atoms with one or seven outer electrons tend to be more reactive, while those with full outer shells are unreactive.

Q201. Describe how Mendeleev arranged elements in his periodic table.
Answer: Mendeleev arranged the known elements in order of increasing atomic weight. However, he didn't rely only on atomic weight—he also placed elements with similar properties into the same columns, or groups. He left gaps where he believed undiscovered elements should go, predicting their properties based on trends he observed. This made his periodic table more useful and accurate than previous attempts.

Q202. Why did Mendeleev leave gaps in his periodic table?
Answer: Mendeleev left gaps in his periodic table because he noticed that some elements did not fit the pattern of properties when arranged strictly by atomic weight. He believed that elements were missing and predicted that they would be discovered later. These gaps allowed elements with similar chemical and physical properties to be grouped together, even if it meant skipping over some atomic weight values.

Q203. How did Mendeleev use predictions to support his periodic table?
Answer: Mendeleev predicted the existence and properties of elements that had not yet been discovered, based on the trends he saw in his periodic table. When these elements were later discovered and matched his predictions closely, it gave strong support to the validity of his table. For example, he predicted the properties of germanium before it was found, and its properties closely matched what he had forecasted.

Q204. What was unusual about the way Mendeleev arranged some elements?
Answer: What was unusual about Mendeleev's arrangement was that he sometimes placed elements out of order by atomic weight. This was because he prioritized grouping elements with similar properties together rather than strictly following atomic weight. For example, he placed iodine after tellurium, even though iodine has a lower atomic weight, because it fit better with other halogens.

Q205. Why did Mendeleev not always follow the order of atomic weight?
Answer: Mendeleev didn't always follow the order of atomic weight because he recognised that chemical properties were more important in grouping elements. When the properties of an element didn't match those in its group, he would change its position even if the atomic weight was lower or higher than nearby elements. This approach helped to maintain consistency in the properties of elements in the same group.

Q206. How did the discovery of isotopes support Mendeleev's decisions?
Answer: The discovery of isotopes helped explain why Mendeleev was correct in not strictly following atomic weights. Isotopes are atoms of the same element with different masses. Since atomic number, not atomic mass, determines chemical behaviour, Mendeleev's decisions made sense. The presence of isotopes meant that atomic weight alone was not a reliable basis for arranging elements, supporting his method.

Q207. What is the importance of testing scientific predictions?
Answer: Testing scientific predictions is important because it allows scientists to check if a new theory is accurate. If predictions made by a theory match experimental results, the theory is supported. If not, it may need to be changed or rejected. This process helps ensure scientific ideas are based on evidence. It also allows science to advance by confirming which theories are reliable and which are not.

Q208. Give an example of how a scientific idea can be supported by experiment.
Answer: One example is how Mendeleev's periodic table was supported by the discovery of new elements like gallium and germanium. He predicted their properties before they were discovered. When these elements were found, their properties matched his predictions, which supported his idea of arranging elements by trends rather than just atomic weight. This experimental confirmation strengthened the validity of his periodic table.

Q209. How can predictions be used to test a new scientific theory?
Answer: A good scientific theory can make clear predictions about what will happen under certain conditions. Scientists can then carry out experiments or observations to see if those predictions come true. If the results match the predictions, the theory is supported. If not, the theory may be flawed. This method is key to developing reliable scientific knowledge that is based on evidence, not guesses.

Q210. What does it mean when a scientific prediction is refuted?
Answer: When a scientific prediction is refuted, it means that the outcome of an experiment or observation does not match what the theory said would happen. This suggests that the theory may be wrong or incomplete. Scientists may then revise or replace the theory with one that better explains the evidence. Refutation helps science move forward and improves our understanding of how things work.

Q211. Why were early periodic tables considered incomplete?
Answer: Early periodic tables were considered incomplete because they did not include all the known elements and could not accurately predict the properties of elements. Some elements were placed in the wrong groups because the order was based only on atomic weight. This led to elements with different properties being grouped together, which made the tables less useful and less reliable compared to later versions like Mendeleev's.

Q212. What made Mendeleev's periodic table more successful than earlier ones?
Answer: Mendeleev's table was more successful because he arranged elements by both atomic weight and chemical properties, and left gaps for undiscovered elements. He also predicted the properties of these elements, and when they were found, they matched his predictions. This showed that his approach worked well and made it easier to understand and classify elements compared to earlier versions.

Q213. How did the discovery of new elements confirm Mendeleev's ideas?
Answer: Mendeleev's ideas were confirmed when new elements like gallium, scandium, and germanium were discovered and had properties very close to what he had predicted. These discoveries filled the gaps he had left and proved that his periodic table could be used to accurately forecast the existence and behaviour of elements not yet known, showing that his method of classification was reliable.

Q214. What is meant by the term "atomic weight"?
Answer: Atomic weight is the average mass of the atoms of an element, taking into account all the naturally occurring isotopes and their relative abundance. It was used in early versions of the periodic table to arrange elements. Today, scientists use atomic number instead because it gives a clearer idea of the element's identity based on the number of protons in its nucleus.

Q215. How does atomic number differ from atomic weight?
Answer: Atomic number is the number of protons in the nucleus of an atom and defines which element it is. Atomic weight, on the other hand, is the average mass of the element's atoms, including all its isotopes. Atomic number is always a whole number, while atomic weight is often a decimal. The periodic table is now arranged by atomic number, which is more reliable than atomic weight.

Q216. Why is the modern periodic table arranged by atomic number?
Answer: The modern periodic table is arranged by atomic number because it reflects the actual number of protons in each atom. This arrangement places elements with similar chemical properties

into the same group. It also avoids the issues caused by relying on atomic weight, especially when elements have isotopes, and ensures that trends in chemical behaviour are more consistent across periods and groups.

Q217. Where are metals found on the periodic table?
Answer: Metals are found on the left side and in the centre of the periodic table. This includes Group 1 and Group 2 metals, as well as the transition metals in the middle block. Metals tend to have few outer electrons and are good conductors of heat and electricity. They are usually solid at room temperature and can form positive ions easily by losing electrons.

Q218. Where are non-metals located in the periodic table?
Answer: Non-metals are found on the right side and towards the top of the periodic table. This includes elements in Groups 5, 6, 7, and 0. Non-metals have more electrons in their outer shell and often gain electrons during reactions. They are poor conductors, and many are gases or brittle solids. Their properties are very different from those of metals.

Q219. Why do metals tend to form positive ions?
Answer: Metals tend to form positive ions because they have only one, two, or three electrons in their outer shell. It is easier for them to lose these electrons and achieve a full outer shell, resulting in a positively charged ion. This process is more energy-efficient than gaining electrons, so metals usually form cations during chemical reactions.

Q220. Why do non-metals usually not form positive ions?
Answer: Non-metals have more electrons in their outer shell, usually five, six, or seven. It would require a lot of energy to remove enough electrons to form a positive ion. Instead, they tend to gain electrons to complete their outer shell, forming negative ions. This process is easier and more stable, which is why non-metals rarely form positive ions.

Q221. How does the position of an element relate to its reactivity?
Answer: The position of an element in the periodic table helps predict its reactivity. Elements in Group 1 are highly reactive metals because they easily lose one electron. Group 7 elements are also reactive because they easily gain one electron. Reactivity in metals increases down the group, while in non-metals, it decreases down the group. The number of outer electrons plays a key role.

Q222. Explain how electron arrangement affects how an element reacts.
Answer: The electron arrangement, especially the number of outer shell electrons, determines how easily an atom can lose, gain, or share electrons. Atoms with nearly full or nearly empty outer shells are more reactive. For example, Group 1 elements react by losing one electron, and Group 7 elements react by gaining one. Atoms with full outer shells, like noble gases, are unreactive.

Q223. Why do elements in the same group react in similar ways?
Answer: Elements in the same group have the same number of electrons in their outer shell. This means they tend to gain or lose the same number of electrons in reactions, leading to similar

chemical behaviour. For example, all Group 1 metals lose one electron and form +1 ions, while all Group 7 non-metals gain one electron to form -1 ions.

Q224. What is the link between group number and number of outer electrons?
Answer: For elements in Groups 1 to 7, the group number tells you how many electrons are in the outer shell. For example, elements in Group 2 have two outer electrons. Group 0 elements have full outer shells—usually 8 electrons—except for helium, which has 2. This link helps predict reactivity and the types of ions that elements form.

Q225. How does atomic structure influence whether an element is a metal or non-metal?
Answer: Atomic structure, especially the number of outer electrons, determines whether an element is a metal or non-metal. Metals have 1 to 3 outer electrons and tend to lose them to form positive ions. Non-metals have 4 to 7 outer electrons and tend to gain electrons to form negative ions. This difference in behaviour defines their physical and chemical properties.

Q226. Why are most elements in the periodic table metals?
Answer: Most elements are metals because the electronic structures that produce metallic behaviour—having only one, two, or three electrons in the outer shell—occur again and again as atomic number rises. After the first energy level is filled, each new period begins with elements that have only a few outer electrons. These atoms lose electrons easily, form positive ions, and bond in giant metallic lattices that give them high electrical and thermal conductivity, malleability, and lustre. Since this low-outer-electron pattern repeats so frequently across periods, the majority of the 118 known elements fall into the metallic category, filling the entire left side and centre of the table while non-metals occupy a narrower region on the right.

Q227. How can you tell if an element is likely to be a metal from its position?
Answer: You can predict an element is a metal if it lies on the left side or in the central block of the periodic table. Elements in Groups 1 and 2 and all the transition blocks (Groups 3–12) are metals, as are many of the lower members of Groups 13 and 14. The position reflects their having few outer electrons, which means they lose electrons easily to form positive ions and exhibit metallic bonding. If an element is below and to the left of the dividing zig-zag line that runs roughly from boron to astatine, it is almost certainly metallic in character.

Q228. What are some physical properties typical of metals?
Answer: Metals usually have a shiny surface when freshly cut, called metallic lustre, and conduct heat and electricity well because free electrons can move through the lattice. They are malleable, meaning they can be hammered into sheets, and ductile, meaning they can be drawn into wires without breaking; these properties arise from layers of atoms that can slide without disrupting metallic bonds. Most are solid at room temperature and have high melting and boiling points due to the strong attraction between positive ions and delocalised electrons within the metallic structure.

Q229. What are some physical properties typical of non-metals?
Answer: Non-metals are generally poor conductors of heat and electricity because they lack free,

mobile electrons. Many are gases at room temperature, such as oxygen and nitrogen; those that are solid, like sulfur or iodine, are often brittle and dull rather than shiny. Their melting and boiling points are usually lower than those of metals because they exist as molecules or have simple atomic structures held together by weak intermolecular forces rather than by a strong metallic lattice, so less energy is needed to separate their particles.

Q230. Describe a key chemical property of metals.
Answer: A key chemical property of metals is their tendency to lose outer-shell electrons to form positive ions (cations) during chemical reactions. This happens readily because their outer electrons are held relatively weakly by the nucleus. Losing electrons allows metal atoms to achieve the stable electronic configuration of the nearest noble gas. As a result, metals react with non-metals to form ionic compounds, react with acids to produce salts and hydrogen gas, and participate in redox reactions where they act as reducing agents.

Q231. Describe a key chemical property of non-metals.
Answer: Non-metals typically gain electrons to achieve a full outer shell, forming negative ions (anions) or sharing electrons through covalent bonding. Because they have more outer electrons and stronger nuclear attraction, adding electrons is easier than losing them. This leads to the formation of covalent molecular substances like water and carbon dioxide, or ionic compounds such as sodium chloride where the non-metal accepts electrons from a metal. Non-metals often act as oxidising agents in redox reactions because they pull electrons away from other species.

Q232. How does the electronic structure of a metal relate to its chemical properties?
Answer: Metals have one to three outer electrons that are relatively weakly held. During reactions these electrons can delocalise, leading to metallic bonding, or be transferred to non-metal atoms to form positive ions. The small number of outer electrons explains high electrical conductivity (electrons are free to move), metallic lustre (electrons reflect light), and reactivity trends such as increasing tendency to lose electrons down a group, because outer electrons become farther from the nucleus and more easily removed.

Q233. How does the electronic structure of a non-metal relate to its chemical properties?
Answer: Non-metals usually possess five to seven outer electrons, giving them a strong pull on additional electrons and allowing them either to gain electrons in ionic bonding or to share electrons in covalent bonding. This high electronegativity leads to the formation of negative ions when reacting with metals, or molecules with other non-metals. Because their outer shells are nearly filled, they do not conduct electricity well (few free electrons) and display varied states (gases or brittle solids) depending on intermolecular forces between their molecules.

Q234. What happens to reactivity as you move down Group 1?
Answer: Reactivity of Group 1 metals increases as you move down the group. Each element has one outer electron, but as atomic radius grows and shielding by inner shells increases, the outer electron is farther from the nucleus and less strongly attracted. This makes it easier to lose the electron, so elements like potassium and cesium react more vigorously with water and oxygen than

lithium or sodium. Therefore the tendency to form +1 ions, release hydrogen, and generate heat and light becomes progressively greater down the column.

Q235. What happens to reactivity as you move up Group 7?
Answer: Reactivity of Group 7 non-metals (the halogens) increases as you move up the group. Halogens need to gain one electron to complete their outer shell. Higher up the group the atoms are smaller, so the outer shell is closer to the nucleus and feels a stronger pull from the positive charge. This stronger attraction allows fluorine and chlorine to attract an extra electron more readily than bromine or iodine. Thus fluorine is the most reactive halogen, while iodine is much less reactive.

Q236. Why are Group 0 elements unreactive?
Answer: Group 0 elements, the noble gases, are unreactive because their atoms have full outer electron shells. This stable configuration means they neither need to gain nor lose electrons, so there is little energetic benefit to forming bonds. The absence of an unfilled outer shell results in extremely low chemical reactivity. Only under extreme conditions, such as very high pressures or with highly electronegative elements like fluorine, do a few noble gases form compounds, and even then only xenon, krypton, or radon participate.

Q237. What type of bonding do metals usually form?
Answer: Metals normally form metallic bonding, where positively charged metal ions are surrounded by a sea of delocalised electrons. These free electrons move through the lattice, holding the ions together by electrostatic attraction. This structure gives metals their characteristic properties such as electrical conductivity, ductility, and high melting points. In reactions with non-metals, metals can also transfer electrons completely, forming ionic bonds and producing metal cations within ionic lattices.

Q238. What type of bonding do non-metals usually form?
Answer: Non-metals typically form covalent bonds by sharing pairs of electrons to complete their outer shells. This sharing produces discrete molecules or giant covalent networks, depending on the element and conditions. When reacting with metals, non-metals may also accept electrons, forming ionic bonds and creating anions that combine with metal cations in crystalline lattices. The precise bonding type depends on the difference in electronegativity between reacting atoms.

Q239. How does the number of shells change as you go down a group?
Answer: As you go down any group in the periodic table, each successive element has one more occupied electron shell than the element above it. This happens because an extra electron shell is needed to accommodate the increasing number of electrons. For instance, lithium has two shells, sodium three, and potassium four. The addition of shells increases atomic radius and decreases the effective nuclear attraction felt by outer electrons, influencing trends such as reactivity and ionisation energy.

Q240. Why do Group 1 metals react more violently down the group?
Answer: Group 1 metals react more violently down the group because the single outer electron is increasingly distant from the nucleus and shielded by more inner shells. The weaker electrostatic

attraction makes it easier to lose that electron, so the metal forms ions faster and releases more energy in reactions. For example, lithium fizzes gently in water, whereas potassium ignites and burns with a lilac flame, and cesium can explode on contact with water, illustrating the rising reactivity.

Q241. How does the size of an atom affect how easily it loses electrons?
Answer: A larger atom has outer electrons that are farther from the nucleus. The increased distance and greater shielding by inner electrons reduce the nuclear attraction on those outer electrons. Consequently, less energy is required to remove them, so the atom loses electrons more readily. This explains why ionisation energy decreases down a group: as atomic radius grows, atoms such as sodium or potassium lose an outer electron more easily than smaller atoms like lithium.

Q242. How does the attraction between nucleus and outer electrons affect reactivity?
Answer: The strength of attraction between the nucleus and outer electrons determines how easily an atom can lose or gain electrons. A weak attraction, caused by large atomic radius or high shielding, allows metals to lose electrons readily, increasing their reactivity. Conversely, a strong attraction in small non-metal atoms enables them to pull in extra electrons easily, making them more reactive oxidising agents. Therefore, variations in nuclear attraction drive the observable trends in reactivity across periods and down groups.

Q243. What do elements in the same period have in common?
Answer: Elements in the same period share the same number of occupied electron shells. As you move across a period, protons and electrons are added to the same outer shell, but no new shells are started. This common shell number explains why atomic radius decreases across a period (nuclear charge increases while shielding stays constant) and why properties change gradually from reactive metals on the left to reactive non-metals and then noble gases on the right.

Q244. What do elements in the same group have in common?
Answer: Elements in the same group have the same number of electrons in their outer shell, giving them similar chemical properties. This common outer-shell electron count leads to consistent valencies, similar types of ions formed, and comparable trends in reactivity. For instance, all Group 2 elements form +2 ions, and all Group 7 elements form -1 ions. Physical properties may vary, but the underlying chemistry is strongly alike because bonding depends on outer electrons.

Q245. Why are transition metals placed in the centre of the periodic table?
Answer: Transition metals occupy the central block (Groups 3–12) because their outer electron configurations involve the gradual filling of d-orbitals. This shared electronic feature leads to characteristic properties such as variable oxidation states, formation of coloured compounds, and catalytic activity. Placing them in a separate block distinguishes them from the main-group (s- and p-block) elements where valence electrons reside only in s and p orbitals, making the table's structure reflect underlying electron configurations.

Q246. How are the properties of transition metals different from Group 1 metals?
Answer: Transition metals are generally harder, stronger, and have much higher melting points than

the soft, low-density Group 1 metals. They show variable oxidation states, form coloured ions, and act as catalysts, while Group 1 metals only form +1 ions and produce colourless compounds. Transition metals are less reactive with water and oxygen, corroding slowly, whereas Group 1 metals react quickly, often violently. These differences arise from their partially filled d-orbitals and stronger metallic bonding.

Q247. How does the periodic table help predict the properties of elements?
Answer: The periodic table organises elements so that position reflects electronic structure. Groups indicate the number of outer electrons, predicting ionic charges and bonding behaviour, while periods show how many shells are occupied, revealing atomic size trends. Moving across or down reveals systematic changes in electronegativity, ionisation energy, and metallic or non-metallic character. By locating an unknown element's atomic number, chemists can anticipate its physical state, likely compounds, and reactivity without direct experimentation.

Q248. Why was arranging elements by atomic number more accurate than by atomic weight?
Answer: Atomic number counts protons, which uniquely defines an element and correlates directly with electron configuration and chemical behaviour. Atomic weight can vary because of isotopes—atoms of the same element with different neutron counts—so some elements appeared out of order when arranged by mass alone. Organising by atomic number eliminates these discrepancies, aligns elements with similar properties into groups more reliably, and explains anomalies Mendeleev observed, such as tellurium and iodine.

Q249. What role did the understanding of subatomic particles play in improving the periodic table?
Answer: Discovering protons, neutrons, and electrons showed that atomic number (proton count) determines chemical identity, while isotopes differ only in neutron number. This clarified why mass-based ordering occasionally failed. Knowledge of electron shells and subshells explained periodicity: repeating chemical patterns arise because similar outer-electron configurations recur as protons are added. Understanding subatomic particles thus provided a theoretical foundation that turned Mendeleev's empirical arrangement into a coherent, predictive scientific model.

Q250. How can the periodic table be used to predict the type of ion an element will form?
Answer: By locating an element's group, you know how many electrons are in its outer shell. Main-group metals in Groups 1, 2, and 13 lose one, two, or three electrons to form +1, +2, or +3 ions. Non-metals in Groups 15, 16, and 17 gain three, two, or one electrons to form -3, -2, or -1 ions. Transition metals often form several positive ions because they can lose differing numbers of d and s electrons. Thus group position guides predictions of ionic charge and bonding behaviour.

Q251. Why are the elements in Group 0 called noble gases?
Answer: They are called noble gases because they are very unreactive. Their atoms have full outer electron shells, making them stable and not needing to react with other elements. This chemical stability is similar to the idea of nobility being above involvement, which is why they are referred to as "noble."

Q252. Why are noble gases generally unreactive?
Answer: Noble gases are generally unreactive because their atoms have complete outer electron shells. This means they do not need to gain, lose, or share electrons to become stable. As a result, they do not easily form compounds or participate in chemical reactions.

Q253. What is the outer electron configuration of helium?
Answer: The outer electron configuration of helium is $1s^2$. Helium has only one electron shell, and that shell is full with two electrons, making it stable and unreactive.

Q254. What is the outer electron configuration of neon?
Answer: The outer electron configuration of neon is $2s^2\ 2p^6$. This means the second shell is fully occupied with eight electrons, which is a stable arrangement and explains neon's lack of reactivity.

Q255. Why does helium only have two electrons in total?
Answer: Helium only has two electrons because its atomic number is 2, meaning it has two protons and therefore two electrons to balance the charge. Its only shell can hold a maximum of two electrons, which is why helium does not have more.

Q256. How does the full outer shell make noble gases stable?
Answer: A full outer shell means that the atom does not need to gain, lose, or share electrons. This complete shell provides chemical stability, so the atom has no tendency to form bonds or react with other atoms, making noble gases very stable and unreactive.

Q257. What is the trend in boiling points of Group 0 elements as you go down the group?
Answer: The boiling points of Group 0 elements increase as you go down the group. Helium has the lowest boiling point, and each element below it has a higher boiling point than the one above due to increased intermolecular forces.

Q258. Why do the boiling points of noble gases increase down the group?
Answer: The boiling points increase down the group because the atoms get larger and have more electrons. This leads to stronger van der Waals forces (intermolecular forces) between the atoms, which require more energy to overcome, hence a higher boiling point.

Q259. How does relative atomic mass affect the boiling point in Group 0?
Answer: As the relative atomic mass increases, the size and number of electrons in the atoms also increase. This results in stronger van der Waals forces between the atoms, which raises the boiling point of the noble gases.

Q260. Predict the boiling point trend for a newly discovered noble gas below radon.
Answer: A noble gas below radon would likely have a higher boiling point than radon. This is because it would have a larger atomic mass and more electrons, leading to stronger intermolecular forces and requiring more energy to boil.

Q261. How do noble gases exist at room temperature?
Answer: Noble gases exist as colourless, odourless gases at room temperature. They are monoatomic, meaning each particle is a single atom. Due to their low boiling points, they remain in the gas state under normal conditions.

Q262. Why do noble gases not form molecules under normal conditions?
Answer: Noble gases do not form molecules because their outer electron shells are full, making them stable. Since they do not need to gain, lose, or share electrons, they remain as single atoms and do not form chemical bonds.

Q263. How can the trend in boiling points help identify an unknown noble gas?
Answer: If you measure the boiling point of an unknown noble gas, you can compare it with the known boiling points of other noble gases. Since boiling point increases with atomic mass, the value can help determine the identity of the gas.

Q264. What would you expect about the density of noble gases down the group?
Answer: The density of noble gases increases as you go down the group. This is because the atoms become heavier (higher atomic mass) and are more tightly packed, even though they are still gases at room temperature.

Q265. Why does the density of noble gases increase as you go down the group?
Answer: The density increases because the atoms have more mass and slightly stronger intermolecular forces, making them heavier and more closely spaced, even in the gas state. The increase in atomic size also contributes to greater density.

Q266. How are the noble gases used in everyday life due to their unreactivity?
Answer: Their unreactivity makes noble gases useful in situations where reactions are unwanted. For example, argon is used in welding to prevent metal from reacting with oxygen. Helium is used in balloons, and neon in signs. They're also used in lighting and insulating windows.

Q267. Why are noble gases used in light bulbs?
Answer: Noble gases are used in light bulbs because they do not react with the hot metal filament. This prevents the filament from burning out quickly, allowing the bulb to last longer. Argon is commonly used for this purpose.

Q268. Why is helium used in balloons instead of hydrogen?
Answer: Helium is used because it is non-flammable and safe. Hydrogen is lighter but highly flammable, which makes it dangerous. Helium provides lift without the risk of fire or explosion.

Q269. Describe the appearance and state of xenon at room temperature.
Answer: Xenon is a colourless, odourless gas at room temperature. It is denser than air and is one of the heavier noble gases. Like others in Group 0, it exists as single atoms and does not form molecules under normal conditions.

Q270. What would be the expected chemical reaction of argon with sodium?
Answer: Argon would not react with sodium under normal conditions. It is a noble gas with a full outer shell, so it is chemically unreactive and does not form compounds easily, even with reactive metals like sodium.

Q271. Why do Group 1 elements have similar chemical properties?
Answer: Group 1 elements all have one electron in their outer shell, which they lose easily in reactions to form +1 ions. This similar electron structure results in similar chemical behaviour, such as reacting with water to form hydrogen gas and an alkaline solution.

Q272. How many electrons are in the outer shell of a Group 1 element?
Answer: Each Group 1 element has one electron in its outer shell. This makes them very reactive as they can lose that one electron easily to achieve a full outer shell, forming a stable ion.

Q273. What is meant by the term "alkali metals"?
Answer: Alkali metals refer to the elements in Group 1 of the periodic table. They are called this because they form alkaline (basic) solutions when they react with water. These metals include lithium, sodium, potassium, and others.

Q274. Describe what happens when lithium reacts with water.
Answer: When lithium reacts with water, it floats and fizzes gently, producing hydrogen gas and lithium hydroxide. The reaction is not violent, and the metal slowly disappears as it forms the alkaline solution.

Q275. Describe what happens when sodium reacts with water.
Answer: When sodium reacts with water, it floats, moves around quickly on the surface, and fizzes more vigorously than lithium. It produces hydrogen gas and sodium hydroxide, which makes the water alkaline. The reaction is more intense and produces heat.

Q276. How is the reaction of potassium with water different from that of lithium?
Answer: Potassium reacts far more vigorously than lithium: it skids rapidly across the water, melts into a silvery sphere, gives off hydrogen so quickly that the gas catches fire, and the surface burns with a lilac-coloured flame; lithium only floats, fizzes gently, and rarely ignites. **Solution:** The key difference is the lower first-ionisation energy of potassium. Its outer electron sits in the fourth shell, is screened by more inner electrons, and feels a weaker pull from the nucleus, so it is lost more readily. Quicker electron loss means faster formation of K^+ and OH^-, releasing more heat. The extra energy melts the metal, accelerates hydrogen evolution, and raises the local temperature above the ignition point of hydrogen, so combustion follows. Lithium's smaller atom holds its outer electron much more strongly, so the reaction proceeds slowly, releases less heat, and seldom reaches the ignition temperature necessary for a flame.

Q277. What gas is produced when Group 1 metals react with water?
Answer: The reaction always produces hydrogen gas. **Solution:** Each alkali-metal atom donates its

single outer electron to the water molecule, creating a metal hydroxide (e.g., NaOH or KOH) and reducing two H_2O molecules to one H_2 molecule. Because metals supply electrons rapidly, hydrogen escapes as bubbles that may ignite if the reaction is vigorous. The presence of hydrogen explains why the mixture may flame or explode and why proper safety measures, such as small pieces and safety screens, are essential in demonstrations.

Q278. What type of solution is formed when Group 1 metals react with water?
Answer: An alkaline solution containing the metal hydroxide is formed. **Solution:** The dissolved product is a strong base such as lithium hydroxide, sodium hydroxide, or potassium hydroxide. These hydroxides fully dissociate into MOH and OH^- ions, giving the solution a high pH (typically above 12). The hydroxide ions turn universal indicator deep purple, feel slippery, and can corrode skin. Because Group 1 hydroxides are highly soluble, the entire mass of metal eventually converts to ions, leaving a clear yet strongly basic solution.

Q279. Write a word equation for the reaction between sodium and water.
Answer: sodium + water → sodium hydroxide + hydrogen. **Solution:** Two atoms of sodium react with two molecules of water to yield two units of sodium hydroxide and one molecule of hydrogen gas. The balanced symbol equation is $2\,Na + 2\,H_2O \rightarrow 2\,NaOH + H_2$. This shows sodium is oxidised (loss of an electron) to Na^+, while water is reduced, releasing H_2. The exothermic nature of the reaction can melt the sodium and may ignite the hydrogen if enough heat accumulates.

Q280. Why does reactivity increase down Group 1?
Answer: Reactivity rises because the outer electron is held less tightly in heavier atoms. **Solution:** Each step down the group adds a new electron shell, so the single valence electron is farther from the nucleus and experiences more shielding by inner electrons. This reduces the effective nuclear attraction, lowers first-ionisation energy, and allows the electron to be removed with less energy. Therefore larger atoms such as potassium or rubidium lose their electron more readily, making their reactions with water or halogens faster and more exothermic than those of lithium.

Q281. How does the atomic structure of potassium explain its higher reactivity?
Answer: Potassium's valence electron is in the 4s sub-shell, far from the nucleus and well shielded by three inner shells. **Solution:** The increase in distance and electron-electron repulsion reduces the effective nuclear charge on that electron, so the ionisation energy drops to about 419 $kJ\,mol^{-1}$, much lower than lithium's 520 $kJ\,mol^{-1}$. Consequently potassium can lose its electron quickly, forming K^+ and releasing more energy in reactions. The high energy release produces heat that melts the metal and ignites hydrogen during the water reaction, illustrating its greater reactivity compared with lithium or sodium.

Q282. Describe the reaction of sodium with chlorine.
Answer: When sodium metal is exposed to chlorine gas it burns with a bright yellow-orange flame, producing white crystals of sodium chloride. **Solution:** The reaction is a vigorous redox process: $2\,Na + Cl_2 \rightarrow 2\,NaCl$. Each sodium atom transfers its single valence electron to a chlorine atom, forming Na^+ and Cl^- ions that assemble into a cubic ionic lattice. The large enthalpy of lattice

formation and the low ionisation energy of sodium release enough heat to sustain the luminous flame. The product, common salt, is stable, crystalline, and unreactive under normal conditions.

Q283. Why is it dangerous to store alkali metals in air?
Answer: They react spontaneously with oxygen and water vapour, risking fire or explosion.
Solution: Even trace moisture on their surfaces can start the formation of metal hydroxides and hydrogen, generating heat that may ignite the hydrogen or the metal itself. Larger pieces can develop oxide layers that crack, exposing fresh metal and causing a chain reaction. Some, like potassium, may ignite in seconds, so exposure can lead to burns, toxic smoke, and violent flames. Hence direct contact with air poses serious hazards.

Q284. What precaution is taken when storing alkali metals?
Answer: They are kept under an inert, oxygen-free medium such as mineral oil or argon. **Solution:** Submerging lithium, sodium, or potassium in paraffin oil excludes air and moisture, preventing oxidation and hydrolysis. For highly reactive metals like rubidium and cesium, sealed ampoules filled with dry argon are preferred. The container is labelled clearly, stored in a cool, dry place, and handled with forceps. These precautions stop accidental ignition and ensure the metal remains in a pure, workable state for laboratory use.

Q285. What is observed when lithium reacts with oxygen?
Answer: Lithium burns with a crimson flame producing a white smoke of lithium oxide. **Solution:** On heating, lithium combines with oxygen in a $4\,Li + O_2 \rightarrow 2\,Li_2O$ reaction. The crimson or deep red colour is characteristic of lithium ions emitting light at about 670 nm. The oxide coats the metal as a white powder that can flake off, exposing more lithium and sustaining combustion. Because the reaction is exothermic, small pieces may glow brightly, but it is less violent than the corresponding reactions of sodium or potassium.

Q286. Why do Group 1 metals form +1 ions?
Answer: Each atom contains a single electron in its outer shell that is easily lost, leaving a stable noble-gas configuration. **Solution:** After losing that one electron, the resulting cation has the electron configuration of the nearest noble gas (e.g., Na^+ matches neon). The energy required to remove the first electron is small, while the second ionisation energy is extremely high due to the stable closed shell now exposed, so only one electron is removed. Consequently all Group 1 reactions produce +1 ions such as Li^+, Na^+, and K^+.

Q287. What is the product when potassium reacts with chlorine?
Answer: The reaction forms potassium chloride, a white crystalline salt. **Solution:** $2\,K + Cl_2 \rightarrow 2\,KCl$. Each potassium atom donates its outer electron to a chlorine atom, yielding K^+ and Cl^- ions. These ions arrange into a face-centred cubic lattice, whose large lattice enthalpy drives the reaction to completion. The product is soluble in water, tastes salty, and is used in salt substitutes for dietary potassium supplementation.

Q288. How does the size of Group 1 atoms affect their reactivity?

Answer: Larger atomic size lowers ionisation energy, so reactivity increases. **Solution:** Going down the group adds electron shells, expanding atomic radius and increasing shielding. The valence electron becomes progressively less tightly bound, so it can be removed by weaker interactions with reactants such as water or halogens. Therefore lithium reacts slowly, sodium reacts faster, and potassium, rubidium, and cesium react explosively. The trend shows a direct link between atomic size, ease of electron loss, and chemical activity.

Q289. Predict the reactivity of rubidium with water based on the group trend.

Answer: Rubidium will react extremely violently, likely exploding on contact with water. **Solution:** Rubidium's outer electron is even farther from the nucleus than potassium's and feels minimal effective nuclear charge, so its first-ionisation energy is only about 403 $kJ\ mol^{-1}$. When dropped into water, electron transfer is almost instantaneous, producing RbOH and H_2 with enormous heat. The heat ignites hydrogen and may shatter the container. Safety protocols require tiny quantities, explosion shields, and remote handling.

Q290. What happens when Group 1 metals react with non-metals?

Answer: They form ionic compounds in which the metal becomes a +1 cation and the non-metal becomes an anion. **Solution:** The metal transfers its single valence electron to the non-metal, creating strong electrostatic attraction that packs the ions into a crystalline lattice. Examples include NaCl, KBr, and Li_2O. These salts have high melting points, conduct electricity when molten or in solution, and are generally white and soluble (except Li_2O, which is sparingly soluble).

Q291. How do halogens exist at room temperature?

Answer: Halogens exist as diatomic molecules: F_2 and Cl_2 are pale gases, Br_2 is a volatile red-brown liquid, and I_2 is a shiny purple solid that sublimes. **Solution:** The X_2 form gives each atom a full octet by sharing a single covalent bond. The physical state varies because the size and polarizability of the molecules increase down the group, strengthening van der Waals forces and raising melting and boiling points. Even iodine molecules are held together only by weak intermolecular forces, so the element can sublime easily.

Q292. Why do halogens form diatomic molecules?

Answer: Each halogen atom has seven valence electrons and needs one more to complete an octet; sharing a pair gives both atoms eight electrons. **Solution:** Forming an X–X single covalent bond lowers the energy compared with isolated atoms. The diatomic structure is stable because the bond strength is moderate (F–F weaker, Cl–Cl stronger) and the resulting molecule has no net charge, making it the default state under standard conditions. The tendency to pair explains why free halogen atoms are transient and highly reactive.

Q293. How many electrons are in the outer shell of a Group 7 element?

Answer: Each Group 7 atom has seven valence electrons. **Solution:** The electronic configuration ends in s^2p^5 for all halogens. Having only one vacancy, they readily gain an electron in ionic reactions to form X^- or share one electron in covalent bonds. This seven-electron valence shell governs their high electronegativity, strong oxidising power, and ability to displace less reactive halogens from compounds.

Q294. Why do halogens have similar chemical properties?
Answer: Their atoms all have the same valence-electron configuration (s^2p^5), so they undergo similar types of reactions. **Solution:** Each halogen seeks one electron to complete its octet, making them powerful oxidising agents. They all form −1 ions with metals, share one electron in covalent molecules, and participate in displacement reactions. Differences in reactivity stem from bond energies and atomic size, but the fundamental chemistry—gaining or sharing one electron—remains consistent across the group.

Q295. Describe what happens when chlorine reacts with iron.
Answer: Heated iron wool glows bright orange-yellow in chlorine gas and forms brown fumes of iron(III) chloride. **Solution:** The reaction $2\,Fe + 3\,Cl_2 \rightarrow 2\,FeCl_3$ is highly exothermic. Chlorine oxidises iron from 0 to +3 while being reduced to Cl^-. The heat generated keeps iron red-hot even after the external flame is removed. The product sublimes as brown-yellow smoke that condenses into a crystalline solid, which is highly soluble in water and forms acidic solutions.

Q296. What kind of compounds do halogens form with metals?
Answer: They form ionic metal halides such as NaCl, $CaCl_2$, and KBr. **Solution:** Metal atoms transfer electrons to halogen atoms, creating oppositely charged ions that assemble into strong lattices. The halides are typically colourless, have high melting points, and conduct electricity when molten. The stoichiometry reflects the metal's valence: Group 1 metals give MX, Group 2 give MX_2, aluminium gives AlX_3, and so on. Halides are widespread in nature and essential in industry and biology.

Q297. Why does reactivity decrease down Group 7?
Answer: Electronegativity and oxidising power decline as atomic size increases. **Solution:** Larger atoms have valence electrons farther from the nucleus and more shielded, so the effective nuclear attraction for an incoming electron is weaker. As a result, the energy released when the atom gains an electron (electron affinity) becomes less negative from fluorine to iodine. Bond strengths in X_2 also drop, so heavier halogens form weaker bonds and are less eager to take electrons, making them progressively less reactive.

Q298. What happens when chlorine is added to a solution of potassium iodide?
Answer: Chlorine displaces iodine, turning the solution brown. **Solution:** $Cl_2 + 2\,KI \rightarrow 2\,KCl + I_2$. Chlorine, being more reactive, oxidises I^- to I_2 while being reduced to Cl^-. The liberated iodine dissolves partly in water giving a brown colour and partly precipitates as tiny dark crystals. Adding starch turns the mixture deep blue-black, confirming iodine's presence. This displacement illustrates the trend that a halogen can displace any other halide below it in the group.

Q299. What is observed when bromine displaces iodine in a solution?
Answer: The orange solution turns dark brown or almost black as iodine forms. **Solution:** When

bromine water is added to potassium iodide, the reaction $Br_2 + 2 I^- \rightarrow 2 Br^- + I_2$ occurs. Bromine is a stronger oxidiser than iodide ion, so it removes an electron to become Br^- while iodide is oxidised to elemental iodine. The iodine's colour change is easily visible; if an organic solvent like hexane is added and shaken, the iodine moves into the organic layer, giving it a violet hue.

Q300. Predict the outcome of a reaction between fluorine and potassium chloride.
Answer: Fluorine will displace chlorine, forming potassium fluoride and releasing chlorine gas.
Solution: 2 KF can be viewed as inadvertent here; actually the reaction is $F_2 + 2 KCl \rightarrow 2 KF + Cl_2$. Fluorine has the highest electronegativity and oxidising power of any element, so it readily removes electrons from Cl^- ions, oxidising them to Cl_2 while being reduced to F^-. The reaction is rapid, exothermic, and potentially violent; it produces white KF powder and green-yellow chlorine gas that must be handled with extreme care.

Q301. Explain why the melting point of iron is much higher than that of sodium.
Answer: Iron has a giant metallic lattice in which each atom releases not only its 4s electron but several 3d electrons, so there are many delocalised electrons moving between closely packed positive ions. The strong electrostatic attraction between these electrons and the high-charge Fe nuclei produces very strong metallic bonds that need a great deal of energy to break. Sodium contributes only one 3s electron per atom, its ions are bigger and farther apart, and the electrostatic forces are much weaker, so its lattice collapses at a much lower temperature. **Solution:** More bonding electrons, closer ion packing and a higher positive charge make iron's metallic bonding far stronger, giving it a melting point near 1538 °C compared with sodium's 98 °C.

Q302. How does the density of copper compare with that of potassium, and what does this show about their atomic structures?
Answer: Copper is nearly nine times denser than potassium because each Cu atom has a much higher atomic mass yet the atoms are packed into a face-centred-cubic lattice with very small spaces between them. Potassium atoms are lighter and their body-centred-cubic lattice has larger gaps, so less mass is contained in the same volume. **Solution:** High atomic mass plus tight packing in transition-metal lattices give copper a high density, while low mass and looser packing in Group 1 metals give potassium a low density that lets it float on water.

Q303. Describe one practical method a student could use to compare the hardness of chromium and lithium.
Answer: Place equal-sized freshly cut rods of chromium and lithium on a firm surface and draw a hardened steel file across each with the same force. Lithium will be cut deeply and smear because its layers slide easily, while chromium will resist scratching and produce only fine metal filings. Record the depth of the grooves and the force needed. **Solution:** A simple scratch or file test shows chromium's strong metallic bonding and closely packed lattice resist deformation, proving it is far harder than lithium, whose single-electron bonds are weak.

Q304. Why is manganese considered stronger than any Group 1 metal when both are subjected to the same mechanical force?

Answer: Manganese atoms are linked by many delocalised electrons from 3d and 4s orbitals, forming strong metallic bonds and a dense lattice that holds atoms firmly. Group 1 metals have only one bonding electron per atom and larger atomic radii, so the bonds are weaker and the ions can slide past each other easily when stress is applied. **Solution:** The larger electron pool and compact structure of manganese oppose dislocation movement, giving high tensile strength, whereas the weak bonding in Group 1 metals makes them soft and easily deformed.

Q305. Discuss how cobalt's reactivity with cold water differs from that of sodium under the same conditions.

Answer: A bright surface of cobalt placed in cold water shows almost no visible change; tiny bubbles may form only after long standing and the metal soon becomes coated with an oxide that stops further attack. Sodium reacts instantly: it floats, darts about, fizzes vigorously, produces hydrogen gas and sodium hydroxide, and enough heat to melt itself. **Solution:** Cobalt's stronger metallic bonds, higher ionisation energy and a protective oxide layer make it essentially unreactive in cold water, while sodium's weak bonding and easy loss of its single valence electron make it violently reactive.

Q306. Outline why nickel does not catch fire in air as quickly as potassium does.

Answer: Nickel forms a thin, adherent oxide layer that blocks oxygen from the fresh metal, and its high melting point means surface heating rarely reaches ignition temperature. Potassium readily loses its single electron to oxygen, releasing heat that melts the metal and exposes more surface, so combustion spreads rapidly. **Solution:** Transition-metal passivation and strong bonding give nickel high fire resistance, whereas the low activation energy and vigorous exothermic oxidation of potassium make it ignite easily in air.

Q307. State the reaction products when copper is heated strongly in oxygen and compare this with the reaction of lithium under the same conditions.

Answer: Heating copper in oxygen forms black copper(II) oxide, CuO, in a slow, surface-only reaction. Heating lithium in oxygen produces white lithium oxide, Li_2O, quickly and exothermically, often accompanied by a bright red flame. **Solution:** The stronger bonds and lower reactivity of copper mean only its surface converts to CuO, while lithium's readiness to donate its electron makes the whole metal react vigorously, forming Li_2O throughout.

Q308. Suggest a reason why iron filings rust in damp air but lithium turns dull grey much faster.

Answer: Iron rusts when water and oxygen together oxidise Fe to hydrated iron(III) oxide, a process that needs time for electrolyte formation. Lithium reacts directly with oxygen and traces of moisture to make Li_2O and $LiOH$ almost immediately; its very high reactivity means a grey layer forms within seconds. **Solution:** Lithium's low ionisation energy and single electron allow instant surface oxidation, whereas iron's higher activation energy slows rusting, so iron filings corrode more slowly than lithium.

Q309. Describe the colour change seen when an iron(II) salt is oxidised to iron(III) and explain the change in ion charge.

Answer: A pale green iron(II) solution turns yellow-brown as Fe^{2+} loses an electron to become Fe^{3+}. The extra positive charge alters ligand–metal interactions and the d–d electron transitions, giving the new colour. **Solution:** Oxidation increases the effective nuclear charge on the d-electrons, changing energy levels and therefore the light absorbed and transmitted, so Fe^{3+} solutions appear brown rather than green.

Q310. Explain why transition metals such as cobalt form more than one stable ion, whereas sodium forms only Na^{+}.

Answer: Cobalt has closely spaced 3d and 4s energy levels, so it can lose different numbers of electrons and still reach relatively stable configurations, giving Co^{2+} and Co^{3+}. Sodium's 3p subshell is far higher in energy than its 3s, so after losing one 3s electron the next electron would be removed from a much lower, stable neon-like shell, requiring very high energy. **Solution:** Small energy gaps between d and s orbitals let transition metals have variable oxidation states; large gaps after the first electron loss restrict Group 1 metals to +1.

Q311. How does the catalytic ability of nickel in hydrogenation reactions reflect typical transition-metal behaviour?

Answer: Nickel provides empty 3d orbitals to adsorb hydrogen and the unsaturated substrate on its surface, weakening their bonds and allowing new H–C bonds to form at lower temperatures and pressures. After the reaction, products leave and the surface is free again. **Solution:** Variable oxidation states and partially filled d orbitals enable temporary adsorption and bond rearrangement without nickel itself being consumed, a hallmark of transition-metal catalysis.

Q312. Chromium forms green Cr^{3+} solutions and yellow CrO_4^{2-} solutions. What does this reveal about variable oxidation states in transition elements?

Answer: The two colours show chromium exists stably in at least two different oxidation states: +3 in Cr^{3+} and +6 in chromate. Each state has different d-electron arrangements and ligand environments, giving distinct colours. **Solution:** Transition elements can lose different numbers of electrons, leading to ions and oxyanions with unique properties and colours; this variable valency is a key transition-metal feature.

Q313. A student adds chlorine gas to an aqueous solution of Fe^{2+} ions. Predict the main observations and justify your answer.

Answer: The pale green solution will turn yellow-brown and may release small bubbles of Cl_2 that dissolve. Chlorine oxidises Fe^{2+} to Fe^{3+}, itself being reduced to Cl^-, so the colour change indicates formation of Fe^{3+}. **Solution:** $Cl_2 + 2\,Fe^{2+} \rightarrow 2\,Cl^- + 2\,Fe^{3+}$; chlorine is a stronger oxidising agent than Fe^{3+}, so it removes an electron from Fe^{2+} and the solution darkens.

Q314. Why does copper not react vigorously with dilute hydrochloric acid while potassium does?

Answer: Copper has a positive electrode potential and strong metallic bonding, so removing electrons to form Cu^{2+} is not favourable in dilute HCl. Potassium has very low ionisation energy; it loses its electron readily to H^+, releasing hydrogen gas violently. **Solution:** Thermodynamics and kinetics favour potassium's oxidation but not copper's, so copper sits inert whereas potassium reacts explosively with dilute acid.

Q315. Outline an experiment to compare the rate at which magnesium ribbon and iron filings react with oxygen at room temperature.
Answer: Weigh equal masses of clean magnesium ribbon and iron filings on separate watch glasses, leave them in open air, and record mass gain at fixed intervals over several days. Plot mass versus time. Magnesium gains mass slowly, forming a thin oxide; iron gains mass more slowly still as it rusts. **Solution:** Differences in mass gain show relative reaction rates, with magnesium oxidising faster than iron under the same conditions.

Q316. Explain why cobalt can act as a catalyst in the Fischer–Tropsch process, whereas lithium cannot.
Answer: Cobalt's surface adsorbs CO and H_2 via its partially filled d orbitals, weakens their bonds, allows C–C coupling and hydrogenation, then releases hydrocarbons. Lithium lacks d orbitals, forms only a strong oxide layer, and cannot bind reactants in the required way. **Solution:** Variable oxidation states and d-orbital availability give cobalt the ability to form transient complexes that lower activation energies; lithium's chemistry is too limited.

Q317. State the observations when nickel metal is placed in cold water for one hour.
Answer: No visible reaction occurs: the nickel stays shiny, no bubbles form, and the water's pH stays neutral. **Solution:** Nickel is effectively inert to cold water because its oxide film and high activation energy prevent electron transfer needed to produce Ni^{2+} and H_2.

Q318. Why does iron(III) chloride appear yellow-brown in solution, whereas sodium chloride is colourless?
Answer: Fe^{3+} has partially filled d orbitals. Ligand field splitting allows d-d transitions that absorb portions of white light, leaving a yellow-brown colour. Na^{+} has no d electrons, so no such transitions occur; all visible light passes through, so the solution is colourless. **Solution:** Colour in many transition-metal ions arises from d-electron excitations, absent in s-block ions like Na^{+}.

Q319. Describe how the strength of a chromium bar can be tested and compared with that of a potassium rod.
Answer: Clamp each bar horizontally and hang increasing masses at the midpoint until each bends permanently. Record the load causing yield. Chromium withstands far greater loads without bending; potassium bends or snaps under a small mass. **Solution:** The higher load before yield reflects stronger metallic bonds and higher tensile strength in chromium.

Q320. Give a reason why manganese can form both Mn^{2+} and MnO_4^- ions, but sodium cannot form Na^{2+} or NaO_2^- ions.
Answer: The 3d and 4s electrons in manganese have similar energies, allowing loss of different numbers to reach several stable states, and the large lattice or solvation energies of its oxyanions stabilise high oxidation states like +7. In sodium the gap after removing the 3s electron is huge, so

removing more is energetically prohibitive. **Solution:** Small energy differences between available orbitals let transition metals show variable valency; large gaps limit Group 1 metals to +1.

Q321. Predict what would happen if a piece of copper is lowered into molten sodium bromide and explain your reasoning.
Answer: No reaction will occur because copper has a lower tendency to be oxidised than bromide has to be reduced; Cu cannot displace Na^+ or Br^-. **Solution:** The electrochemical series shows that Cu^{2+}/Cu is below Na^+/Na but above Br_2/Br^-, so neither oxidation of copper nor reduction of bromide ions is favoured at the temperature of molten NaBr.

Q322. How does the hardness of nickel influence its use in everyday objects compared with the softness of Group 1 metals?
Answer: Nickel's hardness lets it resist scratching, denting and wear, so it is alloyed into coins, stainless steel, and tools. Group 1 metals are so soft they can be cut with a knife and would deform under small loads, making them unsuitable for structural use. **Solution:** Strong metallic bonding and compact lattice give nickel durability needed for practical items, while the weak bonds in Group 1 metals restrict them to chemical applications.

Q323. Suggest why copper pipes resist corrosion better than iron pipes in ordinary tap water.
Answer: Copper forms a thin, adherent layer of Cu_2O that seals the surface and stops further attack, and its electrode potential is higher than that of dissolved oxygen in neutral water, so oxidation is slow. Iron forms porous rust that flakes away, exposing fresh metal to water and oxygen. **Solution:** A protective oxide film and favourable electrochemical potential make copper corrosion-resistant, whereas unstable rust layers keep iron corroding.

Q324. Describe a simple test that distinguishes between aqueous Fe^{2+} and Fe^{3+} ions using sodium hydroxide solution.
Answer: Add a few drops of NaOH. Fe^{2+} gives a green gelatinous precipitate of $Fe(OH)_2$ that quickly turns brown on standing; Fe^{3+} gives an immediate brown precipitate of $Fe(OH)_3$. **Solution:** The colour difference on formation and on exposure to air allows easy identification of the oxidation state.

Q325. Explain how ligand exchange reactions in transition metals give rise to colour changes, using aqueous copper complexes as an example.
Answer: In water, copper(II) exists mainly as pale blue $[Cu(H_2O)_6]^{2+}$. Adding concentrated ammonia replaces four water ligands with NH_3 to make deep blue $[Cu(NH_3)_4(H_2O)_2]^{2+}$. The change in ligand field alters d-orbital splitting, so the wavelength of light absorbed shifts; we see a new colour. **Solution:** Ligand substitution changes the electronic environment of the metal ion, modifying d-d transition energies and producing observable colour changes typical of transition-metal chemistry.

Q326. Compare the behaviour of chromium and sodium when both are exposed to dry chlorine gas at room temperature.
Answer: Chromium reacts slowly with chlorine gas at room temperature, forming chromium(III) chloride, $CrCl_3$. The reaction is not vigorous due to chromium's resistance to oxidation. In contrast,

sodium reacts very rapidly and exothermically with chlorine, forming sodium chloride with bright flames.

Equation: $2Cr + 3Cl_2 \rightarrow 2CrCl_3$

Equation: $2Na + Cl_2 \rightarrow 2NaCl$

Q327. Why do transition metals generally have higher tensile strength than Group 1 metals?

Answer: Transition metals have more delocalised electrons and stronger metallic bonding due to d-orbital involvement. Their closely packed atoms resist deformation, giving them high tensile strength. Group 1 metals, in contrast, have larger atomic radii, weaker bonds, and only one delocalised electron, making them soft and easy to deform under stress.

Q328. Outline why the density trend across Cr, Mn, Fe, Co, Ni, Cu does not mirror the trend in Group 1 metals.

Answer: In transition metals, atomic size changes very little across the period while atomic mass increases, leading to higher densities. Their atoms are packed tightly in the metallic lattice. In contrast, Group 1 elements have increasing atomic size and loosely packed structures as you go down the group, resulting in lower densities.

Q329. Describe the effect of adding ammonia to a solution of copper(II) sulfate and explain the colour changes observed.

Answer: When ammonia is added to $CuSO_4$ solution, a pale blue precipitate of $Cu(OH)_2$ forms first. On excess ammonia, the precipitate dissolves to form a deep blue solution due to the formation of $[Cu(NH_3)_4(H_2O)_2]^{2+}$ complex ions. This ligand exchange changes the d-orbital splitting, causing a different colour to be absorbed and reflected.

Equation: $[Cu(H_2O)_6]^{2+} + 4NH_3 \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+} + 4H_2O$

Q330. Explain why potassium must be stored under oil but copper can be left in air without serious risk.

Answer: Potassium reacts rapidly with oxygen and moisture in air, forming potassium oxide and releasing heat, which may cause it to ignite. Copper, however, reacts very slowly and forms a protective layer of copper oxide that prevents further reaction.

Equation for potassium: $4K + O_2 \rightarrow 2K_2O$

Equation for copper (slow): $2Cu + O_2 \rightarrow 2CuO$

Q331. Discuss how variable oxidation states in iron contribute to its role in biological systems such as haemoglobin.

Answer: Iron in haemoglobin exists mainly as Fe^{2+}, which can bind oxygen reversibly. The ability of iron to switch between Fe^{2+} and Fe^{3+} is crucial for oxygen transport and release in the body. This redox flexibility allows it to act as a carrier of oxygen in blood and plays a key role in respiration processes.

Q332. A student heats cobalt metal in steam. Predict the main chemical change and write a chemical equation.

Answer: When cobalt is heated in steam, it reacts to form cobalt(II) oxide and hydrogen gas. This is a redox reaction where cobalt is oxidised and water is reduced.
Equation: $Co + H_2O (g) \rightarrow CoO + H_2$

Q333. Why are transition-metal ions often coloured, while most Group 1 metal ions are colourless in solution?
Answer: Transition-metal ions have partially filled d orbitals. When light passes through their solutions, certain wavelengths are absorbed as electrons move between split d-orbitals. The remaining wavelengths give the solution its colour. Group 1 ions have no available d–d transitions, so they do not absorb visible light and appear colourless.

Q334. Evaluate the suitability of manganese dioxide as a catalyst in the decomposition of hydrogen peroxide compared with a Group 1 metal compound.
Answer: Manganese dioxide (MnO_2) is an excellent catalyst for decomposing H_2O_2 because it uses variable oxidation states to lower the activation energy. Group 1 metal compounds lack this flexibility and are not effective catalysts.
Equation: $2H_2O_2 \rightarrow 2H_2O + O_2$ (catalysed by MnO_2)

Q335. State two properties of nickel that make it useful in making stainless steel.
Answer: Nickel is corrosion-resistant and adds strength to steel. It helps form a stable, protective oxide layer on the steel surface and improves the hardness of the alloy, making it suitable for structural, medical, and household applications.

Q336. Compare the lattice energies of sodium chloride and iron(III) oxide and relate these to their melting points.
Answer: Iron(III) oxide has much higher lattice energy due to its Fe^{3+} and O^{2-} ions with greater charges, resulting in stronger electrostatic attractions. This contributes to its much higher melting point compared to NaCl, which has lower charges (Na^+ and Cl^-).
NaCl equation (melts at ~800°C): $Na^+ + Cl^-$
Fe_2O_3 (melts at ~1565°C): $2Fe^{3+} + 3O^{2-}$

Q337. Suggest why the electrical conductivity of copper is higher than that of lithium even though both are metals.
Answer: Copper has more delocalised electrons from both s and d orbitals, allowing more charge carriers to move freely. Its atomic structure is also more compact, reducing resistance. Lithium has only one delocalised electron and a less dense structure, resulting in lower conductivity.

Q338. Describe how the reactivity of iron with dilute acids differs from that of potassium and explain why.
Answer: Iron reacts slowly with dilute acids, producing hydrogen gas gradually. Potassium reacts explosively with dilute acids, releasing hydrogen rapidly and generating heat.
Equation (iron): $Fe + 2HCl \rightarrow FeCl_2 + H_2$
Equation (potassium): $2K + 2HCl \rightarrow 2KCl + H_2$

Q339. Explain the term "complex ion" and give an example involving chromium(III).
Answer: A complex ion consists of a central metal ion surrounded by ligands that donate lone pairs of electrons. In chromium(III), an example is $[Cr(H_2O)_6]^{3+}$, where six water molecules act as ligands forming coordinate bonds with the Cr^{3+} ion. This bonding affects colour, solubility, and reactivity.

Q340. Predict the products when iron reacts with bromine vapour and justify your answer in terms of oxidation states.
Answer: Iron reacts with bromine to form iron(III) bromide. Iron is oxidised from 0 to +3 and bromine is reduced from 0 to –1.
Equation: $2Fe + 3Br_2 \rightarrow 2FeBr_3$

Q341. Why does adding chloride ions to a cobalt(II) solution change its colour from pink to blue?
Answer: In aqueous solution, Co^{2+} exists as $[Co(H_2O)_6]^{2+}$, which is pink. When Cl^- ions are added, they replace water ligands to form $[CoCl_4]^{2-}$, which is blue. This ligand exchange changes the d-orbital splitting and alters the light absorbed and reflected.

Q342. Discuss the significance of high density in copper when it is used for electrical wiring.
Answer: Copper's high density means more atoms and free electrons per unit volume, allowing greater current flow with minimal resistance. This makes copper wires compact yet highly efficient, which is ideal for household and industrial wiring systems.

Q343. Outline an investigation to compare the catalytic actions of cobalt and nickel in the hydrogenation of vegetable oils.
Answer: Heat equal amounts of unsaturated oil with hydrogen in presence of cobalt in one setup and nickel in another. Keep pressure, temperature, and time constant. After the reaction, measure the iodine value or melting point to determine the degree of saturation and compare catalytic efficiency.

Q344. Explain why transition metals show magnetic properties, using iron as an example.
Answer: Iron has unpaired electrons in its 3d orbitals, which create magnetic moments. These moments align under an external magnetic field, giving iron ferromagnetic properties. Group 1 metals lack unpaired d electrons, so they do not show strong magnetic behaviour.

Q345. Describe what is meant by disproportionation and give an example involving manganese compounds.
Answer: Disproportionation is a redox reaction where the same element is both oxidised and reduced. In acidic solution, MnO_4^{2-} can disproportionate to MnO_4^- and Mn^{2+}.
Equation: $3MnO_4^{2-} + 4H^+ \rightarrow 2MnO_4^- + Mn^{2+} + 2H_2O$

Q346. Compare the behaviour of chromium and lithium in a flame test and account for any colour observed.
Answer: Lithium gives a bright crimson-red flame due to excitation of its outer electron. Chromium

usually shows no visible colour in a standard flame test because transitions in d-orbitals are less energetic or fall outside the visible range.

Q347. Why is copper sulphate solution blue, and how would the colour change if the copper(II) ions were reduced to copper(I)?

Answer: Copper(II) ions in $[Cu(H_2O)_6]^{2+}$ absorb light in the red region, reflecting blue light. On reduction to Cu^+, the d orbitals are filled, so no d–d transitions occur, making the solution colourless or very pale.

Equation: $2Cu^{2+} + e^- \rightarrow 2Cu^+$

Q348. State why sodium ions do not act as catalysts in industrial chemical processes, but iron ions do.

Answer: Sodium only forms Na^+ and cannot change oxidation state. Catalysis often involves temporary changes in oxidation states, which iron can do as it exists in both Fe^{2+} and Fe^{3+} forms. This makes iron useful in redox catalysis, unlike sodium.

Q349. Predict how the hardness of cobalt affects its machinability compared with the softness of sodium.

Answer: Cobalt's hardness makes it wear-resistant but harder to cut or shape using tools. Sodium is soft and easy to cut but unsuitable for structural purposes due to its low mechanical strength. Cobalt is harder to machine but ideal for durable applications.

Q350. Explain why adding a small amount of carbon can further strengthen iron, whereas similar alloying with potassium is not feasible.

Answer: In iron, small carbon atoms fit into interstitial spaces of the metallic lattice, restricting atomic movement and increasing strength. Potassium's soft and reactive nature, along with its low melting point, makes it unsuitable for alloying with carbon, as it forms unstable or reactive compounds.