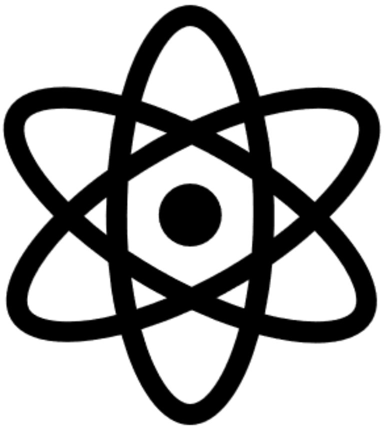
Atomic structure and periodic table



L1 Atoms, elements, compounds – revision

Name ________________________________

Class ________________________________

Teacher ________________________________

Atoms and Elements

Everything around you is made up of tiny particles called atoms. Imagine atoms as the building blocks of everything in the universe. An atom is the smallest part of an element that can exist. Think of elements as different types of atoms. Each element is a substance that cannot be broken down into simpler substances. For example, gold is an element, and its smallest part is a gold atom.

Atoms of each element are represented by a unique chemical symbol, usually one or two letters. For example, the chemical symbol for oxygen is O, and the symbol for sodium is Na. There are about 100 different elements, and you can find all of them listed in the periodic table. The periodic table is like a chart that organizes all the elements by their properties.

Compounds and Chemical Reactions

When elements combine, they form compounds. A compound is a substance made from two or more different elements that are chemically combined. This means the atoms of the elements join in specific ways. For example, water is a compound made from hydrogen and oxygen. Its chemical formula is H_2O, meaning it has two hydrogen atoms and one oxygen atom.

Chemical reactions are processes where substances change to form new substances. During a chemical reaction, atoms rearrange to create something different, often with a noticeable energy change, such as heat or light. Compounds can only be broken down into their elements through chemical reactions.

Chemical reactions can be shown using word equations or symbol equations. For example, the reaction between hydrogen and oxygen to form water can be written as:

Word equation: Hydrogen + Oxygen → Water

Symbol equation: $2H_2 + O_2 \rightarrow 2H_2O$

The Periodic Table and Chemical Symbols

In your studies, you will be given a periodic table, which you can use to find the names and symbols of elements. You should be familiar with the first 20 elements, as well as elements in Groups 1 and 7. For example, the first element is hydrogen (H), and the second is helium (He).

You'll also learn to name compounds. For example, NaCl is sodium chloride, commonly known as table salt. Additionally, you will practice writing word equations and formulae for chemical reactions. A balanced chemical equation has the same number of each type of atom on both sides of the equation.

Mixtures

Unlike compounds, mixtures are combinations of two or more elements or compounds that are not chemically combined. This means each substance in a mixture keeps its own properties. For example, a salad is a mixture because you can still see and separate the different ingredients like lettuce, tomatoes, and cucumbers.

Mixtures can be separated using physical processes, which don't involve chemical reactions. Here are some methods:

1. **Filtration**: Used to separate a solid from a liquid. For example, separating sand from water.
2. **Crystallization**: Used to form solid crystals from a solution. For example, making rock candy from sugar water.
3. **Simple Distillation**: Used to separate a liquid from a solution. For example, distilling water to remove impurities.
4. **Fractional Distillation**: Used to separate a mixture of liquids with different boiling points. For example, separating crude oil into gasoline and other products.
5. **Chromatography**: Used to separate different substances in a mixture. For example, separating the different dyes in ink.

When given information about a mixture, you should be able to suggest the best method for separating it. For instance, if you have a mixture of sand and salt, you could dissolve the salt in water, filter out the sand, and then evaporate the water to get the salt back.

Understanding these concepts will help you recognize how different substances interact and change, which is fundamental in chemistry. Keep practicing writing equations and using the periodic table, and soon you'll have a solid grasp of these essential scientific ideas.

Independent practice

Comprehension Questions

1. What is an atom, and why is it important in the study of chemistry?
2. How are elements different from compounds?
3. Describe what happens during a chemical reaction.
4. What is the purpose of the periodic table, and how are elements represented on it?

Understanding Questions

1. Give the chemical symbols for the following elements: Oxygen, Sodium, and Helium.
2. What is the chemical formula for water, and what elements does it contain?
3. Explain how a compound can be separated into its elements.
4. Describe two methods that can be used to separate mixtures.
5. Why do mixtures retain the properties of their individual components?
6. Write a word equation for the reaction between hydrogen and chlorine to form hydrogen chloride.

Finish the Sentence Questions

1. The periodic table is useful for organizing elements because...
2. Water can be separated into hydrogen and oxygen, but...
3. A mixture of sand and salt can be separated by dissolving the salt in water and...

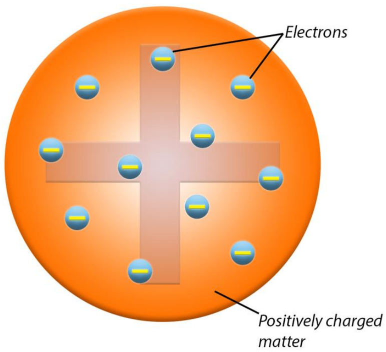
L2 History of the discovery of atomic structure

Understanding the development of the model of the atom is like uncovering the secrets of a mystery. Scientists have been curious about what everything around us is made of for hundreds of years, and their discoveries have shaped the way we understand the world today. Let's go on a journey through time to see how their ideas about the atom evolved.

Early Ideas: Indivisible Spheres

A long time ago, before anyone knew about electrons or nuclei, scientists thought that atoms were the smallest pieces of matter. They believed that atoms were tiny, solid spheres that couldn't be divided into anything smaller. This idea made sense because they couldn't see anything smaller than atoms with the technology they had at the time.

The Discovery of the Electron: The Plum Pudding Model

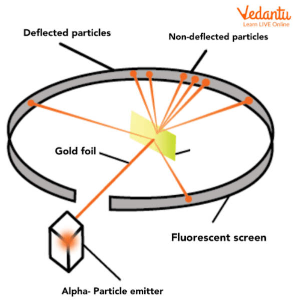


In 1897, a scientist named J.J. Thomson discovered something surprising: the electron. Electrons are tiny particles with a negative charge, and Thomson realized that they must be a part of atoms. But how could tiny spheres contain even tinier particles?

Thomson proposed a new model for the atom called the "plum pudding model." Imagine a pudding (or a muffin) with bits of fruit scattered throughout it. Thomson suggested that an atom was like that pudding, where the positive charge was spread out like the pudding, and the electrons were scattered within it like the bits of fruit. This was a big step forward, but it wasn't quite right.

The Alpha Particle Scattering Experiment: The Nuclear Model

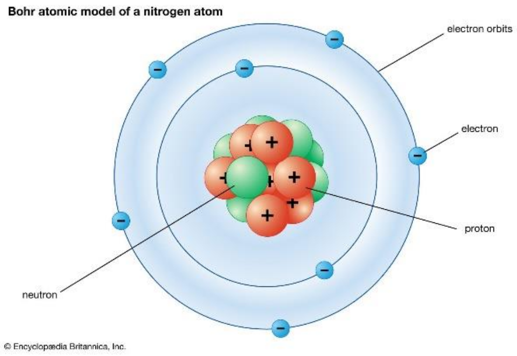
A few years later, in 1909, Ernest Rutherford conducted an experiment that changed everything. He and his team fired tiny particles called alpha particles at a very thin sheet of gold foil. They expected the particles to go straight through, with just a few deflecting slightly. To their surprise, while most alpha particles did pass through, some were deflected at large angles, and a few even bounced back.



Key Observations and Conclusions:

1. **Most alpha particles passed straight through the gold foil:**
   - **Conclusion:** Atoms are mostly empty space. If the atom was solid as in the plum pudding model, most particles would have been deflected. The fact that most passed through suggested that there was nothing in their way most of the time.
2. **Some alpha particles were deflected at large angles:**
   - **Conclusion:** The deflections indicated that there was a concentrated positive charge within the atom. For the alpha particles (which are positively charged) to be repelled, they had to encounter a strong positive force. This led Rutherford to conclude that this positive charge must be concentrated in a small area, which he called the nucleus.

3. **A few alpha particles bounced straight back:**
   - **Conclusion:** The mass of the atom must be concentrated in this small, dense nucleus. Only something very dense and massive could cause the alpha particles to bounce back. This was the definitive evidence that most of the mass of the atom was located in the nucleus.



Bohr's Model: Electrons in Specific Orbits

Niels Bohr, a scientist working in the early 20th century, improved upon Rutherford's model. Bohr suggested that electrons don't just orbit the nucleus randomly. Instead, they travel in specific paths or orbits, much like planets around the sun. Each orbit corresponds to a certain energy level. Bohr's theoretical calculations matched experimental observations, and his model helped explain how atoms absorb and emit light.

Discovery of the Proton: Subdividing the Nucleus

As scientists continued to study the nucleus, they realized that the positive charge of the nucleus wasn't just one single blob of positive stuff. Instead, it could be broken down into smaller particles, each with the same positive charge. These particles were named protons. This discovery helped explain why elements had different properties and masses.

Chadwick's Contribution: The Neutron

About 20 years after the nucleus was first proposed, James Chadwick made another crucial discovery in 1932. He found evidence for a new type of particle in the nucleus, which had no charge. He called these particles neutrons. Neutrons, along with protons, make up the nucleus of an atom. The discovery of neutrons helped explain why atoms of the same element could have different masses, leading to the concept of isotopes.

Why the Atomic Model Changed

The atomic model changed over time because new experimental evidence was discovered that didn't fit the old models. When scientists find new evidence, they have to update their models to match what they observe. This is a fundamental part of science: always being ready to change our understanding based on new information.

Comparing Models

- **Plum Pudding Model:** The atom is a ball of positive charge with negative electrons scattered throughout, like raisins in a pudding.
- **Nuclear Model:** The atom has a small, dense, positively charged nucleus at the center, with electrons orbiting around it.

By understanding how the model of the atom has evolved, we see how scientific knowledge builds over time. Each new discovery helps us get a clearer picture of the world around us.

Independent practice

Comprehension Questions

1. What was the initial idea about the structure of atoms before the discovery of the electron?
2. Describe the plum pudding model of the atom proposed by J.J. Thomson.
3. What were the key observations from Rutherford's alpha particle scattering experiment?
4. How did Niels Bohr modify Rutherford's nuclear model of the atom?

Understanding Questions

1. Why did most alpha particles pass through the gold foil in Rutherford's experiment?
2. Explain why some alpha particles were deflected at large angles in the alpha particle scattering experiment.
3. What conclusion did Rutherford draw from the observation that a few alpha particles bounced back?
4. How did the discovery of protons and neutrons further refine the model of the atom?
5. What is the significance of Bohr's model in explaining atomic behaviour, especially in terms of energy levels?
6. Why is it important for scientific models to change based on new experimental evidence?

Find the Mistake Questions

1. In the plum pudding model, electrons are thought to orbit the nucleus at specific distances.
    - Find the mistake and correct it.
2. Rutherford's experiment showed that the mass of the atom is evenly distributed throughout the atom.
    - Find the mistake and correct it.
3. James Chadwick discovered that the nucleus is made up of electrons and neutrons.
    - Find the mistake and correct it.

L3 Understanding atoms

Atoms are the fundamental building blocks of everything around us. Imagine them as tiny, invisible Lego bricks that make up everything in the universe, from the air we breathe to the water we drink and even to our own bodies.

Structure of an Atom

An atom is made up of three main particles: protons, neutrons, and electrons. These particles are incredibly small, much smaller than anything we can see with our eyes. Here's a closer look at each one:

1. **Protons**: Protons are positively charged particles that are found in the center of an atom, called the nucleus. They are like the + sign in a battery, which carries a positive charge. Each atom has a specific number of protons, which determines what type of element it is. For example, if an atom has 6 protons, it is carbon.

2. **Neutrons**: Neutrons are also found in the nucleus, alongside protons. They don't have an electrical charge, so they are neutral. Neutrons act like glue that holds the nucleus together.

3. **Electrons**: Electrons are tiny, negatively charged particles that zoom around the nucleus in a cloud-like area called shells or energy levels. They are like little planets buzzing around a sun. Electrons are involved in chemical reactions and how atoms bond together to form molecules.

Electrical Charge and Balance

In an atom, the number of protons is always equal to the number of electrons. Protons are positive, and electrons are negative, so their charges cancel each other out. This means that atoms overall have no charge; they are neutral.

Size and Mass

Atoms are incredibly small. Imagine a dot so tiny you can't even see it with the most powerful microscope. That dot is about the size of an atom! Typically, atoms have a radius of about 0.1 nanometers (nm), which is one billionth of a meter.

The nucleus, where protons and neutrons are packed together, is even smaller. It's like having a tiny ball inside a large stadium; the nucleus is about 100,000 times smaller than the atom itself! Most of the atom's mass, or weight, is concentrated in this tiny nucleus.

Mass of Particles

When we talk about the mass of particles within an atom, here's how they compare:

- **Protons**: Protons are relatively heavy compared to electrons. They weigh about 1 atomic mass unit (amu).

- **Neutrons**: Neutrons are also about the same weight as protons, roughly 1 amu.

- **Electrons**: Electrons are much lighter than protons and neutrons. In fact, their mass is so tiny that it's almost negligible when compared to protons and neutrons.

Bringing It All Together: Atomic Number and Mass Number

- **Atomic Number**: This is the number of protons in an atom. It's like the atom's ID number. For example, hydrogen atoms always have 1 proton, so their atomic number is 1.

- **Mass Number**: This is the total number of protons and neutrons in an atom. It tells us how heavy the atom is. For instance, carbon usually has 6 protons and 6 neutrons, so its mass number is 12.

Atoms are amazing in how they combine and form everything we see around us. By understanding their tiny parts — protons, neutrons, and electrons — we can begin to understand how the entire universe is built.

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Independent practice

Comprehension Questions:

1. How do protons, neutrons, and electrons contribute to the overall structure of an atom?
2. Why are atoms considered neutral despite containing charged particles like protons and electrons?
3. What is the relationship between an atom's atomic number and its identity as a specific element?
4. How does the size of an atom compare to the size of its nucleus?

Understanding Questions:

1. Explain why the mass of an atom is primarily concentrated in its nucleus.
2. Describe the role of electrons in determining how atoms interact with each other to form molecules.
3. Compare the relative masses of protons, neutrons, and electrons in an atom.
4. How does the number of protons in an atom relate to its position on the periodic table?

Sentence Completion Questions:

1. Atoms are electrically neutral ___________.
2. The nucleus of an atom contains most of its mass ___________.
3. Electrons are much lighter than protons and neutrons ___________.

[Use "because" "hence," or "whereas" to complete the sentences.]

L4 Isotopes

Atoms are the fundamental units of matter, and each type of atom is classified as an element. Inside an atom, there are three key components: protons, which carry a positive charge; neutrons, which are neutral; and electrons, which carry a negative charge and orbit the nucleus.

Now, let's delve into isotopes. Isotopes are different versions of the same element, characterized by having the same number of protons (which defines the element) but differing numbers of neutrons. For example:

1. **Hydrogen (H)**:
   - Hydrogen-1 (protium): 1 proton, 0 neutrons
   - Hydrogen-2 (deuterium): 1 proton, 1 neutron
   - Hydrogen-3 (tritium): 1 proton, 2 neutrons (radioactive)
2. **Carbon (C)**:
   - Carbon-12: 6 protons, 6 neutrons (stable, most common)
   - Carbon-13: 6 protons, 7 neutrons (stable)
   - Carbon-14: 6 protons, 8 neutrons (radioactive, used in carbon dating)
3. **Oxygen (O)**:
   - Oxygen-16: 8 protons, 8 neutrons (stable, most common)
   - Oxygen-17: 8 protons, 9 neutrons (stable)
   - Oxygen-18: 8 protons, 10 neutrons (stable)
4. **Uranium (U)**:
   - Uranium-235: 92 protons, 143 neutrons (used in nuclear reactors and weapons)
   - Uranium-238: 92 protons, 146 neutrons (most common isotope of uranium)

Isotopes can exhibit different properties and behaviors due to their varying numbers of neutrons. Some isotopes are stable and do not change over time, while others are unstable (radioactive) and decay into other elements over time. For example, Uranium-235 is used in nuclear reactors because it can undergo controlled nuclear fission reactions, releasing energy.

In fields like archaeology, isotopes are used to determine the age of ancient artifacts. For instance, Carbon-14 dating relies on measuring the decay of Carbon-14 to estimate how long ago organic material was alive. In medicine, radioactive isotopes are used in imaging techniques and cancer treatments.

Understanding isotopes is crucial because it allows scientists to study the properties of elements in more detail and apply their knowledge in various practical applications.

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Independent practice

Comprehension Questions:

1. What defines an isotope of an element?
2. How do isotopes of an element differ from each other?
3. Give an example of two isotopes of carbon and their differences in neutron numbers.
4. Name two fields where isotopes are used and explain their applications.

Understanding Questions:

1. Why do isotopes of an element have different numbers of neutrons?
2. How does the number of protons determine the identity of an element?
3. Explain the difference between a stable isotope and a radioactive isotope.
4. Can you describe how isotopes help scientists understand the age of ancient artifacts?

Spot and Correct Mistakes Questions:

1. Identify the mistake in the statement: "Isotopes are atoms that have different numbers of protons and electrons."
    - Correct the mistake and explain the correct components that vary in isotopes.
2. Find the error in the statement: "Carbon-13 has 7 protons and 6 neutrons."
    - Correct the mistake and provide the correct information.
3. What's wrong with the statement: "Oxygen-17 is more common than Oxygen-16."
    - Correct the mistake and explain the correct abundance of oxygen isotopes.
4. Locate the error in the sentence: "Isotopes are only used in chemistry."
    - Correct the mistake and explain other fields where isotopes find applications.

L5 Numbers of protons, neutrons and electrons

Atoms and Their Structure

Imagine atoms as tiny building blocks that make up everything around us, like the bricks in a Lego set. They are incredibly small and are made up of even smaller particles. The main parts of an atom are called protons, neutrons, and electrons.

- **Protons** are positively charged particles found in the center of the atom, called the nucleus.
- **Neutrons** are neutral particles, meaning they have no charge, also found in the nucleus.
- **Electrons** are negatively charged particles that zoom around the nucleus in a kind of cloud.

The number of protons in an atom determines what kind of element it is. For example, if an atom has 6 protons, it's carbon; if it has 8 protons, it's oxygen.

Calculating Protons, Neutrons, and Electrons

To find out how many protons, neutrons, and electrons are in an atom:

- The **atomic number** tells us the number of protons.
- The **mass number** tells us the total number of protons and neutrons combined.

(Mass number) (Atomic number) $^{23}_{11}Na$

For example, if an atom has an atomic number of 6 and a mass number of 12:

- Protons = 6 (because atomic number = number of protons)
- Neutrons = 12 - 6 = 6 (because mass number - atomic number = number of neutrons)
- Electrons = 6 (because in a neutral atom, the number of electrons equals the number of protons)

Relative Atomic Mass

Atoms of the same element can have different numbers of neutrons; these are called isotopes. The relative atomic mass of an element takes into account the average mass of all its isotopes, weighted by how common each isotope is in nature.

For example, carbon has two common isotopes: carbon-12 and carbon-13. Carbon-12 is more common, so it contributes more to the average atomic mass of carbon.

Calculating Relative Atomic Mass

To calculate the relative atomic mass (sometimes called atomic weight):

- Look at the percentage abundance of each isotope.
- Multiply each isotope's mass number by its percentage abundance (in decimal form).
- Add these values together to find the relative atomic mass.

For instance, if carbon-12 has an abundance of 98.89% and carbon-13 has an abundance of 1.11%:

- Relative atomic mass of carbon = (12 * 0.9889) + (13 * 0.0111) = 12.01

This means that on average, a carbon atom has an atomic mass of about 12.01.

Independent practice

Comprehension Questions:

1. What are the three main parts of an atom, and where are they located?
2. How does the number of protons determine the identity of an atom?
3. What are isotopes, and how do they contribute to the concept of relative atomic mass?
4. Explain how the relative atomic mass of an element is calculated using the percentages of its isotopes.

Calculation Questions:

1. How many neutrons are there in an atom with an atomic number of 8 and a mass number of 16?
2. If an atom has 20 protons and a mass number of 40, how many neutrons does it have?
3. Determine the number of neutrons in an atom with an atomic number of 13 and a mass number of 27.
4. Calculate the number of neutrons in an atom with an atomic number of 17 and a mass number of 35.
5. Given the isotopes carbon-12 (98.89% abundance) and carbon-13 (1.11% abundance), calculate the relative atomic mass of carbon.
6. If an element has two isotopes with masses 35 amu (80% abundance) and 37 amu (20% abundance), what is its relative atomic mass?
7. Determine the relative atomic mass of an element with two isotopes: one with a mass of 64 amu (60% abundance) and another with a mass of 66 amu (40% abundance).
8. Calculate the relative atomic mass of an element that has three isotopes with masses 32 amu (20% abundance), 34 amu (50% abundance), and 36 amu (30% abundance).
9. Given the isotopes oxygen-16 (99.76% abundance) and oxygen-18 (0.20% abundance), calculate the relative atomic mass of oxygen.
10. An element has three isotopes with masses 25 amu (10% abundance), 26 amu (20% abundance), and 27 amu (70% abundance). What is its relative atomic mass?

L6 electron configuration

Atoms are the building blocks of everything around us, including you and me! Imagine atoms as tiny, incredibly busy solar systems. Just like how the sun and planets interact, atoms have a central part called the nucleus, and swirling around it are electrons, which are like tiny planets.

What is Electronic Structure?

The electronic structure of an atom tells us about the arrangement of these electrons. Think of it as a map that shows where each electron lives around the nucleus. Each electron has a special home called an energy level or a shell. These shells are like different floors in a skyscraper, but instead of being stacked on top of each other, they surround the nucleus.

How Electrons Occupy Energy Levels

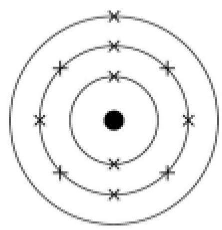
Now, electrons are quite particular about where they live. They always choose the lowest available energy level first before moving to higher ones. It's a bit like you choosing the comfiest chair in the room before considering sitting on the couch.

In the center of an atom, there's a tiny, dense nucleus made up of protons and neutrons. Protons have a positive charge, neutrons have no charge, and together they form the nucleus. This nucleus is surrounded by clouds of electrons whizzing around it.

Visualizing Electronic Structure

We can describe the electronic structure of an atom in two ways: with numbers or with diagrams. Let's take sodium as an example. Sodium is a metal you might have seen in your kitchen or in science class.

- **Numbers:** The electronic structure of sodium is 2,8,1. This means:
  - There are 2 electrons in the first energy level (closest to the nucleus).
  - There are 8 electrons in the second energy level.
  - And finally, there is 1 electron in the third energy level (the farthest one out).
- **Diagram:** If we draw a diagram of sodium's electronic structure, it might look like this:



  - The first energy level (closest to the nucleus) has 2 electrons.
  - The second energy level has 8 electrons.
  - The third energy level has 1 electron.

This diagram helps us visualize where each electron lives in the atom. Each energy level can hold only a certain number of electrons before they move to the next level.

Worked example

Oxygen has 8 electrons. The atomic number of oxygen tells us it has 8 protons in its nucleus, which means it also has 8 electrons to balance the charge (since atoms are neutral overall, having equal numbers of protons and electrons).

Filling Electrons in Oxygen:

1. First Energy Level:
    - This energy level can hold a maximum of 2 electrons.
    - Oxygen's first 2 electrons will go into this shell.
2. Second Energy Level:
    - This energy level can hold a maximum of 8 electrons.
    - After filling the first shell, 6 more electrons need a home.

Energy Level 2: 6 electrons

3. Third Energy Level:
    - This energy level can also hold a maximum of 8 electrons.
    - The remaining 2 electrons will go into this shell.

Energy Level 3: 2 electrons

Summary:

In the case of oxygen (O), which has 8 electrons:

- The first energy level fills up with 2 electrons.
- The second energy level fills up with 6 electrons.
- The third energy level fills up with the remaining 2 electrons.

This gives us the electronic structure of oxygen as 2,6. This means:

- 2 electrons in the first energy level.
- 6 electrons in the second energy level.

Answering Questions about Energy Levels or Shells

When we talk about the electronic structure of atoms, we might use the terms "energy levels" or "shells" interchangeably. Imagine these terms like different ways to describe the same thing. You can choose whichever one helps you understand better!

Why It Matters

Understanding the electronic structure of atoms helps scientists and chemists predict how atoms will behave. It's like knowing the layout of a house—you can figure out where everything is and how it might interact with other houses nearby.

In summary, the electronic structure of an atom tells us where electrons live around the nucleus. They occupy energy levels or shells, starting with the lowest available one first. This structure can be represented by numbers (like 2,8,1 for sodium) or diagrams that show how electrons are arranged. Whether we talk about energy levels or shells, we're describing the same thing: the cozy homes of electrons within atoms.

Independent practice

1. How many electrons does oxygen (O) have in total? (Hint: Use the AQA periodic table to find the atomic number.)
2. What is the maximum number of electrons that can fit in the first energy level of an atom?
3. Describe the electron configuration of sodium (Na) using both numbers and a diagram.
4. Why do electrons occupy the lowest energy levels first in an atom?

Understanding Questions:

1. Explain why it's essential for electrons to fill the lowest energy levels before moving to higher ones.
2. Compare and contrast the terms "energy levels" and "shells" when discussing the electronic structure of an atom.

Practice Questions (Drawing Electron Configurations of Elements):

1. Draw the electron configuration of chlorine (Cl) using both numbers and a diagram. (Hint: Use the AQA periodic table to find the atomic number.)
2. Illustrate the electron configuration of carbon (C). (Hint: Use the AQA periodic table to find the atomic number.)
3. Show the electron configuration for magnesium (Mg) using both numbers and a diagram. (Hint: Use the AQA periodic table to find the atomic number.)
4. Draw the electron configuration of fluorine (F). (Hint: Use the AQA periodic table to find the atomic number.)
5. Illustrate the electron configuration of helium (He). (Hint: Use the AQA periodic table to find the atomic number.)
6. Show the electron configuration for lithium (Li) using both numbers and a diagram. (Hint: Use the AQA periodic table to find the atomic number.)

L7 The development of the periodic table

Imagine a puzzle that scientists worked on for many years — a puzzle made of tiny building blocks called elements. Each element is like a unique Lego piece with its own special properties. Scientists wanted to arrange these elements in a logical way so they could understand them better.

Early Attempts at Classification

Long ago, before scientists knew about protons, neutrons, and electrons, they began trying to organize elements based on their atomic weights. They thought that elements should be lined up from lightest to heaviest. This was a good start, but it didn't always work perfectly. Some elements didn't fit well into this order because their properties didn't match their atomic weights.

Mendeleev's Insightful Solution

Then along came a scientist named Dmitri Mendeleev. He looked at the elements and realized there was a better way to organize them. Mendeleev created a table where elements were arranged not just by weight, but by their properties too. He left spaces in his table for elements that hadn't been discovered yet, predicting what their properties might be based on the patterns he saw.

Discoveries that Filled the Gaps

Something amazing happened — scientists found the missing elements that Mendeleev predicted! They fit right into the spots he left in his table. This showed everyone that Mendeleev's idea was really smart and useful.

Why Scientists Found Mendeleev's Periodic Table Reliable

Scientists found Mendeleev's periodic table reliable for several reasons:

1. **Predictive Power:** Mendeleev's table wasn't just a list of elements — it predicted the existence and properties of elements that hadn't been discovered yet. When these elements were later found and matched the predictions, it showed that Mendeleev's method was accurate.
2. **Organized Patterns:** Mendeleev's arrangement of elements showed clear patterns in their properties. Elements with similar properties were grouped together, which helped scientists see relationships and make connections between different elements.
3. **Adjustable Order:** Unlike earlier tables based solely on atomic weight, Mendeleev's table allowed for adjustments based on properties. Sometimes elements were placed in a different order to better fit their properties, not just their weight.
4. **Scientific Acceptance:** Mendeleev's periodic table was widely accepted by scientists because it provided a framework that made sense of the known elements at the time and made reliable predictions about undiscovered ones.

Understanding Isotopes

Later, scientists learned about isotopes. Isotopes are versions of an element that have the same number of protons but different numbers of neutrons. This discovery explained why sometimes the order based on atomic weights didn't match the order of properties. Isotopes helped scientists understand why some elements seemed out of place in the old tables.

Steps in Developing the Periodic Table

To understand how the periodic table got to where it is today, we can follow these important steps:

1. **Early Classification Attempts:** Scientists first tried to organize elements by atomic weight, but this didn't always fit their properties.

2. **Mendeleev's Breakthrough:** Dmitri Mendeleev improved the periodic table by organizing elements by both atomic weight and their properties. He also predicted the existence of undiscovered elements.
3. **Discovery of New Elements:** Scientists found the elements Mendeleev predicted, which filled the gaps in his periodic table and confirmed his ideas.
4. **Understanding Isotopes:** Learning about isotopes helped scientists understand why some elements didn't follow the exact order of atomic weights. Isotopes explained these differences.

Why It's Important

The periodic table is important because it helps scientists understand how elements work and how they might behave in different situations. It's like a big map that shows us all the elements and how they are related to each other. By studying the periodic table, scientists can make new discoveries and create new materials that can help people in many ways.

In conclusion, the periodic table started with simple ideas about atomic weight but became much more complex and useful thanks to scientists like Mendeleev. It's a great example of how science grows and improves over time, solving puzzles and making the world easier to understand.

Understanding the periodic table is like solving a big mystery where the clues are all the different elements around us. It's a puzzle that scientists are always working on, trying to learn more about the tiny building blocks that make up everything in the universe.

Independent Practice

Comprehension Questions:

1. How did scientists originally attempt to classify elements before the discovery of protons, neutrons, and electrons?
2. What were some limitations of early periodic tables that organized elements based solely on atomic weights?
3. How did Mendeleev improve upon earlier periodic tables?
4. Why were the discoveries of elements that filled Mendeleev's predicted gaps significant?

Understanding Questions:

1. Explain why Mendeleev's periodic table was considered more reliable than earlier classifications based on atomic weight.
2. How did the discovery of isotopes contribute to a better understanding of the periodic table's organization?
3. Why is it important for scientists to organize elements based on both atomic weight and properties rather than just one factor?
4. Compare and contrast Mendeleev's approach to organizing the periodic table with earlier attempts. What were the key differences?

Sentence Completion and Extension:

Finish the sentence and add two more sentence to make a mini paragraph:

1. Sentence: "The periodic table is like a map because..."
2. Sentence: "Discoveries that filled the gaps in Mendeleev's periodic table..."

L8 The modern periodic table

Understanding the Periodic Table and Elements

The periodic table arranges elements by their atomic number, which indicates the number of protons in their nucleus. Elements with similar properties are grouped together in vertical columns called groups or families. This organization is "periodic" because it shows repeating patterns of properties across rows, known as periods.

Atomic Structure and Groups: Elements within the same group share the same number of electrons in their outer shell. Electrons are tiny particles that orbit around the nucleus of an atom. This shared electron configuration gives elements in a group similar chemical behaviour. For instance, elements in the same group often react similarly with other substances.

Predicting Reactions: The position of an element in the periodic table helps scientists predict how it will react with other elements. Metals, located on the left side and toward the bottom of the table, tend to lose electrons easily and form positive ions. Non-metals, found on the right side and toward the top, typically gain or share electrons and are less likely to form positive ions.

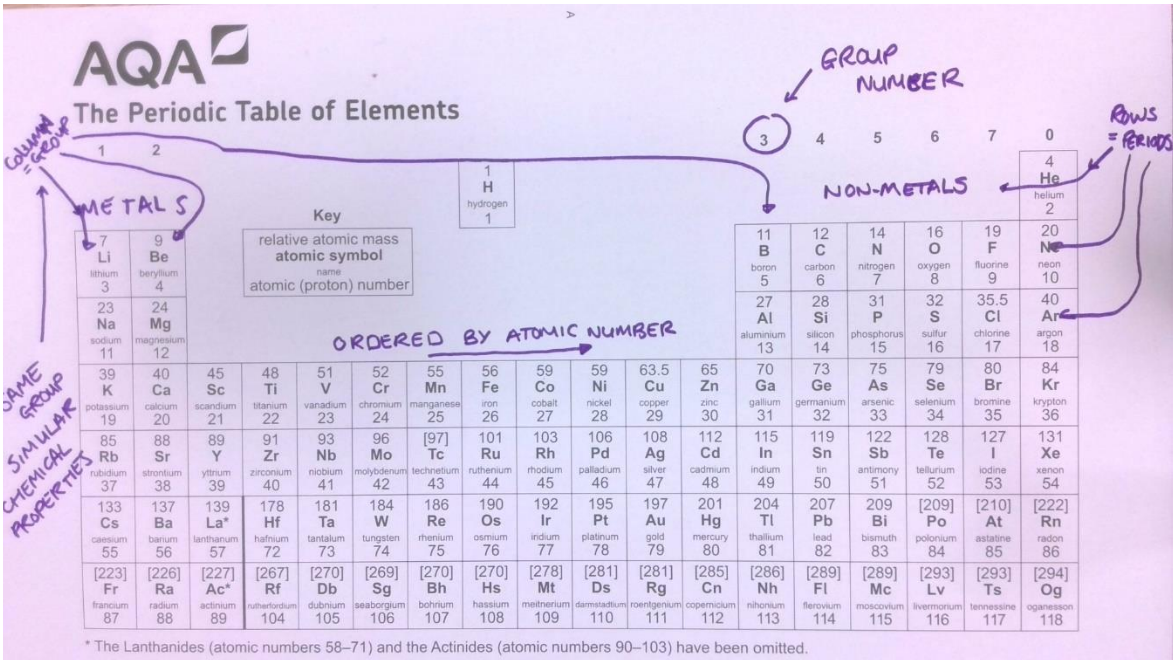
Metals and Non-Metals

Identification: Metals and non-metals exhibit distinct characteristics. Metals are generally shiny, good conductors of heat and electricity, and can be shaped without breaking (malleable) or drawn into wires (ductile). Non-metals can be gases like oxygen and solids like sulphur, often lack lustre, and are poor conductors of heat and electricity.

Atomic Structure and Properties: The atomic structure, including the number of protons, neutrons, and electrons, determines whether an element is categorized as a metal or non-metal. Metals typically have fewer electrons in their outer shell compared to non-metals, which influences their reactivity and ability to form compounds.

Reactions and Atomic Structure: Elements positioned close to each other in the periodic table tend to have similar electron arrangements and therefore exhibit comparable chemical behaviours. This understanding allows scientists to predict which elements will react together and how they will interact in chemical reactions.

In conclusion, the periodic table serves as a crucial tool for understanding how elements behave based on their atomic structure. By organizing elements into groups and periods, it reveals patterns in their properties and facilitates predictions about their chemical reactions. Understanding the distinctions between metals and non-metals helps us comprehend their roles in various natural and technological processes.



Independent Practice

Comprehension Questions:

1. What is a group in the periodic table?
2. How are elements arranged within a group?
3. Define a period in the periodic table.
4. Where are metals typically found on the periodic table? Why?
5. Explain why elements in the same group have similar chemical properties.
6. What does the atomic number of an element indicate?

Understanding Questions:

1. How does the position of an element in the periodic table relate to its atomic structure?
2. Why are non-metals located predominantly on the right side of the periodic table?
3. Compare and contrast the characteristics of metals and non-metals based on their properties.
4. Predict how an element located in Group 17 of the periodic table would likely react with an element from Group 1.

Finding and Correcting Mistakes:

1. "Elements in a period of the periodic table have the same number of electrons in their outer shell." True or False? If false, correct the statement.
2. "All elements on the left side of the periodic table are non-metals." True or False? If false, correct the statement.
3. "Metals are found primarily in periods towards the top of the periodic table." True or False? If false, correct the statement.
4. "Elements in the same group of the periodic table have different numbers of protons." True or False? If false, correct the statement.

L9 The groups

Understanding Group 1: The Alkali Metals

What are Alkali Metals?

Group 1 of the periodic table contains the alkali metals. These elements have some special properties because they all have one electron in their outer shell. This group includes elements like lithium (Li), sodium (Na), and potassium (K).

Reactivity of Alkali Metals

One key thing to know about alkali metals is that their reactivity increases as you go down the group. This means lithium is the least reactive, while potassium is more reactive. The single electron in the outer shell is easily lost during chemical reactions, making these metals very reactive.

Reactions with Oxygen, Chlorine, and Water

- **Oxygen:** When alkali metals react with oxygen, they form metal oxides. For example, when sodium reacts with oxygen, it forms sodium oxide (Na_2O). This reaction can be quite vigorous, especially with potassium, which can catch fire.
- **Chlorine:** Alkali metals react with chlorine to form metal chlorides. For instance, sodium reacts with chlorine to produce sodium chloride (NaCl), commonly known as table salt. These reactions are also very exothermic, meaning they release a lot of heat.
- **Water:** The reaction with water is particularly dramatic. When sodium reacts with water, it forms sodium hydroxide (NaOH) and hydrogen gas (H_2). The reaction releases so much heat that the hydrogen gas can catch fire. For potassium, the reaction is even more explosive.

Why are Alkali Metals so Reactive?

The reactivity of alkali metals depends on the single electron in their outer shell. This election is held relatively loosely because it is far from the nucleus, especially in the heavier alkali metals. This makes it easier for the electron to be lost, leading to a reaction.

Predicting Properties

Based on the trend down the group, we can predict that caesium (Cs) and francium (Fr), which are below potassium, would be even more reactive. This trend helps chemists understand how these elements might behave in different situations.

Understanding Group 7: The Halogens

What are Halogens?

Group 7 contains the halogens, which are non-metals. This group includes elements like chlorine (Cl), bromine (Br), and iodine (I). Halogens have seven electrons in their outer shell, which makes them highly reactive as they seek to gain one more electron to achieve a full outer shell.

Reactivity of Halogens

Unlike alkali metals, the reactivity of halogens decreases as you go down the group. This means chlorine is more reactive than bromine, which is more reactive than iodine. The desire to gain one electron is strongest in the smaller atoms at the top of the group.

Nature of Compounds Formed

- **With Metals:** When halogens react with metals, they form ionic compounds. For example, when chlorine reacts with sodium, it forms sodium chloride (NaCl). In this compound, chlorine gains an electron to form a chloride ion (Cl^-), and sodium loses an electron to form a sodium ion (Na^+).
- **With Non-Metals:** When halogens react with non-metals, they form covalent compounds. For instance, chlorine can react with hydrogen to form hydrogen chloride (HCl), where they share electrons to achieve full outer shells.

Trends in Physical Properties

As you move down the group, the halogens have higher relative molecular masses, and their melting and boiling points increase. This means that chlorine, which is a gas at room temperature, has a lower boiling point than bromine, which is a liquid, and iodine, which is a solid.

Why do Halogens React in this Way?

The properties of halogens depend on the seven electrons in their outer shell. The smaller the atom, the stronger it attracts additional electrons to complete its shell. As the atoms get bigger down the group, the attraction for extra electrons gets weaker, making the elements less reactive.

Displacement Reactions

A more reactive halogen can displace a less reactive halogen from a compound in an aqueous solution. For example, if you add chlorine to a solution of potassium bromide (KBr), the chlorine will displace the bromine, forming potassium chloride (KCl) and bromine gas (Br_2).

Predicting Properties

Based on these trends, we can predict that astatine (At), which is below iodine, would have an even higher melting and boiling point but would be less reactive than iodine. This understanding helps scientists anticipate how these elements will behave in different chemical reactions.

Summary

In summary, Group 1 alkali metals are very reactive because of their single outer electron, with reactivity increasing down the group. Group 7 halogens are also very reactive due to their seven outer electrons, but their reactivity decreases as you move down the group. Understanding these patterns helps us predict how these elements will react in different situations.

Independent practice

Group 1: Alkali Metals

1. What happens to the reactivity of alkali metals as you move down Group 1 of the periodic table?
2. Describe the reaction of sodium with water. What are the products of this reaction?
3. Why are alkali metals very reactive? Explain in terms of their electron configuration.

Group 7: Halogens

1. How does the reactivity of halogens change as you go down Group 7 of the periodic table?
2. What type of compounds do halogens form when they react with metals? Give an example.
3. Why do halogens have similar reactions? Explain in terms of their electron configuration.

Understanding Questions

1. Explain why potassium is more reactive than sodium.
2. Predict what would happen if bromine was added to a solution of potassium iodide. Explain your reasoning.
3. Compare the physical states of chlorine, bromine, and iodine at room temperature. How do these states relate to their positions in Group 7?
4. Describe how the properties of alkali metals and halogens depend on the number of electrons in their outer shells.

Find the Mistake Questions

1. Identify and correct the mistake in the following paragraph:
   - "Lithium, sodium, and potassium are all part of Group 1, known as the halogens. They have two electrons in their outer shell, which makes them very reactive."
2. Identify and correct the mistake in the following paragraph:
   - "Chlorine, bromine, and iodine are alkali metals found in Group 7. Their reactivity increases as you move down the group, with iodine being the most reactive."
3. Identify and correct the mistake in the following paragraph:
   - "When sodium reacts with chlorine, it forms sodium hydroxide and chlorine gas. This reaction is typical for alkali metals reacting with non-metals."