

Chapter 3: Atomic Structure

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Short Questions (Exercise)

(i) Distinguish between shell and sub-shell.

- **Shell:** Represents the main energy level of electrons (K, L, M, N).
- **Sub-shell:** Subdivisions within a shell based on orbitals (s, p, d, f).

(ii) Why is an atom electrically neutral?

An atom is electrically neutral because the number of protons (positive charge) equals the number of electrons (negative charge), balancing the overall charge.

(iii) How many sub-shells are there in the N shell?

The N shell contains **4 sub-shells: s, p, d, and f**.

(iv) Give notation for sub-shells of M shell.

The sub-shells of the M shell are represented as **3s, 3p, and 3d**.

(v) List the sub-shells of M shell in order of increasing energy.

The order of increasing energy for M shell sub-shells is: **3s < 3p < 3d**.

(vi) Can you identify an atom without knowing the number of neutrons in it?

Yes, an atom can be identified by its atomic number, which represents the number of protons (and electrons). The number of neutrons does not affect identification.

3. The electronic configurations listed are incorrect. Explain what mistake has been made in each and write correct electronic configurations.

- **Mistakes:**

The configurations do not follow the Aufbau principle, which states that electrons fill orbitals in order of increasing energy. Higher energy orbitals (like 3d) should not be filled before lower ones (like 3s or 3p).

- **Correct Configurations:**

- $x=1s^2, 2s^2, 2p^6, 3p^1$ $x = 1s^2, 2s^2, 2p^6, 3p^1$
 - $y=1s^2, 2s^2, 2p^1$ $y = 1s^2, 2s^2, 2p^1$
 - $z=1s^2, 2s^2, 2p^6, 3s^1$ $z = 1s^2, 2s^2, 2p^6, 3s^1$
-

4. Which orbital in each of the following pairs is lower in energy?

1. **2s or 2p: 2s**

- The 2s orbital is closer to the nucleus and has lower energy compared to the 2p orbital.

2. **3p or 3s: 3s**

- The 3s orbital has a lower principal quantum number and energy than the 3p orbital.

3. **3s or 4s: 3s**

- The 3s orbital is from a lower energy shell, making it lower in energy compared to the 4s orbital.
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5. Draw Bohr's Model for the following atoms indicating the location for electron, protons, and neutrons.

(a) Potassium (Atomic No. 19, Mass No. 39):

- **Protons:** 19
- **Neutrons:** 20
- **Electrons:** 19
Bohr's model shows:
- 2 electrons in the K shell.
- 8 electrons in the L shell.
- 8 electrons in the M shell.
- 1 electron in the N shell.

(b) Silicon (Atomic No. 14, Mass No. 28):

- **Protons:** 14
- **Neutrons:** 14

- **Electrons: 14**
Bohr's model shows:

- 2 electrons in the K shell.
- 8 electrons in the L shell.
- 4 electrons in the M shell.

(c) Argon (Atomic No. 18, Mass No. 39):

- **Protons: 18**
- **Neutrons: 21**
- **Electrons: 18**
Bohr's model shows:
- 2 electrons in the K shell.
- 8 electrons in the L shell.
- 8 electrons in the M shell.

6. Write the electronic configuration for the following elements:

1. **(a) Oxygen (Atomic No. 8):** $1s^2, 2s^2, 2p^4$
 - Oxygen has a total of 8 electrons distributed across the 1s, 2s, and 2p orbitals.
 2. **(b) Neon (Atomic No. 10):** $1s^2, 2s^2, 2p^6$
 - Neon has 10 electrons, with a completely filled outer shell making it a noble gas.
 3. **(c) Magnesium (Atomic No. 12):** $1s^2, 2s^2, 2p^6, 3s^2$
 - Magnesium has 12 electrons, with 2 electrons in the outermost shell.
 4. **(d) Aluminum (Atomic No. 13):** $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$
 - Aluminum has 13 electrons, with one electron in the 3p sub-shell.
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7. State the importance and uses of isotopes in various fields of life.

- **Medicine:** Isotopes like Cobalt-60 are used in cancer treatment, and iodine-131 is used in diagnosing and treating thyroid diseases.
- **Archaeology:** Carbon-14 isotopes are used in radiocarbon dating to determine the age of ancient artifacts.
- **Energy:** Uranium-235 is a critical isotope used in nuclear reactors for generating electricity.
- **Industry:** Isotopes like Americium-241 are used in smoke detectors.
- **Biology:** Isotopes are used in tracer studies to track metabolic pathways.

8. The atomic number of an element is 23, and its mass number is 56.

(a) How many protons and electrons does this atom have?

The element has **23 protons** (atomic number) and **23 electrons** because the number of electrons equals the number of protons in a neutral atom.

(b) How many neutrons does this atom have?

Neutrons = Mass number - Atomic number = $56 - 23 = 33$

9. The atomic symbol of aluminum is written as Al_{13}^{27} . What information do you get from it?

- **13:** The atomic number, which indicates the number of protons and electrons in the aluminum atom.
- **27:** The mass number, which is the sum of protons and neutrons.
- **Neutrons:** Calculated as $27 - 13 = 14$.

This notation shows that aluminum has 13 protons, 13 electrons, and 14 neutrons.

Extera Short Questions (Topic Wise)

3.1: Atomic Models

1. **What was Dalton's atomic model?**

Dalton proposed that matter is made up of small, indivisible particles called atoms. He stated that atoms of the same element are identical, and chemical reactions involve the rearrangement of atoms.

Example: Water (H_2O) consists of hydrogen and oxygen atoms rearranged in a specific ratio.

2. **How did Rutherford's model describe the atom?**

Rutherford described the atom as having a dense, positively charged nucleus with electrons orbiting around it in empty space. His gold foil experiment showed that most alpha particles passed through but some were deflected, proving the existence of the nucleus.

Example: The nucleus of gold (Au) deflected the alpha particles in his experiment.

3. **What was the main flaw in Bohr's atomic model?**

Bohr's model suggested that electrons orbit in fixed paths around the nucleus. However, it failed to explain the behavior of multi-electron atoms and the spectral lines of larger elements.

4. **What is the modern atomic model?**

The quantum mechanical model describes electrons as existing in regions of space called orbitals, where there is a high probability of finding them, instead of fixed paths.

Example: The electron cloud model for hydrogen shows a high probability of electrons around the nucleus.

3.2: Subatomic Particles

1. **What are the three main subatomic particles?**

Atoms consist of protons (positively charged), neutrons (neutral), and electrons (negatively charged).

2. **Who discovered the electron?**

J.J. Thomson discovered the electron through his cathode ray experiment, where a beam of negatively charged particles was deflected by electric and magnetic fields.

Example: Electrons are represented as e^- in chemical reactions like $\text{Na} \rightarrow \text{Na}^+ + e^-$.

3. **What is the charge of a proton?**

A proton has a positive charge (+1). It is found in the nucleus and determines the atomic number of an element.

4. **Which subatomic particle has no charge?**

Neutrons have no charge. They help stabilize the nucleus by reducing repulsion between protons.

Example: The nucleus of carbon-12 contains 6 protons and 6 neutrons.

3.3: Protons and Atomic Number

1. **What is the atomic number?**

The atomic number is the number of protons in the nucleus of an atom. It uniquely identifies an element.

Example: Carbon has an atomic number of 6, meaning it has 6 protons.

2. **Why is the atomic number important?**

The atomic number determines the element's identity and its position on the periodic table.

Example: Hydrogen, with an atomic number of 1, is the first element on the periodic table.

3. **Can the atomic number of an element change?**

No, the atomic number is fixed for each element. However, nuclear reactions can change the element by altering the number of protons.

4. **Give an example of an atomic number.**

Oxygen has an atomic number of 8, meaning it has 8 protons in its nucleus.

3.4: Relative Atomic Mass and Atomic Mass Unit

1. **What is relative atomic mass?**

Relative atomic mass is the weighted average mass of an element's isotopes, measured relative to 1/12th the mass of a carbon-12 atom.

Example: Chlorine has a relative atomic mass of 35.5 due to the abundance of isotopes Cl-35 and Cl-37.

2. **What is the atomic mass unit (amu)?**

The atomic mass unit is defined as 1/12th the mass of a carbon-12 atom. It is a standard unit for measuring atomic masses.

3. **Why is relative atomic mass not a whole number?**

It accounts for the natural abundance of isotopes with different masses.

Example: Boron has isotopes with masses of 10 and 11, making its average atomic mass approximately 10.8.

4. **Give an example of relative atomic mass.**

Sodium's relative atomic mass is 23, as most of its atoms have this mass.

3.5: Isotopes

1. **What are isotopes?**

Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons.

Example: Carbon-12 and Carbon-14 are isotopes of carbon, with 6 protons but 6 and 8 neutrons, respectively.

2. **Give an example of isotopes.**

Hydrogen has three isotopes: Protium (^1H), Deuterium (^2H), and Tritium (^3H).

3. **How do isotopes differ in properties?**

Isotopes have identical chemical properties but differ in physical properties such as mass and

radioactivity.

Example: Carbon-14 is radioactive, while Carbon-12 is stable.

4. **Why are isotopes important?**

Isotopes are used in various fields such as medical imaging, radiocarbon dating, and nuclear energy.

Example: Cobalt-60 is used in cancer treatment.

3.6: Cations and Anions

1. **What is a cation?**

A cation is a positively charged ion formed when an atom loses electrons.

Example: Na^+ is a cation formed when sodium loses one electron.

2. **What is an anion?**

An anion is a negatively charged ion formed when an atom gains electrons.

Example: Cl^- is an anion formed when chlorine gains one electron.

3. **Give an example of a cation.**

Magnesium ion (Mg^{2+}) is a cation formed when magnesium loses two electrons.

4. **Give an example of an anion.**

Sulfide ion (S^{2-}) is a common anion formed when sulfur gains two electrons.

3.7: Electronic Configuration

1. **What is electronic configuration?**

Electronic configuration shows the arrangement of electrons in an atom's orbitals.

Example: Carbon's configuration is $1s^2 2s^2 2p^2$.

2. **What is the electronic configuration of hydrogen?**

The electronic configuration of hydrogen is $1s^1$, with one electron in its first shell.

3. **How are electrons arranged in orbitals?**

Electrons fill orbitals in increasing energy levels, following the Aufbau principle.

Example: Oxygen's configuration is $1s^2 2s^2 2p^4$, where electrons occupy lower-energy orbitals first.

4. **What is the electronic configuration of neon?**

Neon's configuration is $1s^2 2s^2 2p^6$, with a completely filled second energy level.

Extera Long Questions (Topic Wise)

1. How have atomic models evolved over time, and what are the key contributions of each model?

Dalton's Atomic Model

- Proposed in the early 19th century, John Dalton described atoms as indivisible particles that make up matter.
- Atoms of the same element are identical, and compounds are combinations of different atoms in fixed ratios.
- Example:** Water (H_2O) is a combination of hydrogen and oxygen atoms in a 2:1 ratio.

Thomson's Plum Pudding Model

- J.J. Thomson discovered the electron and proposed that atoms are made of a positively charged "pudding" with negatively charged electrons embedded within.
- Example:** Like raisins in a cake, electrons are scattered within a positively charged sphere.

Rutherford's Nuclear Model

- Ernest Rutherford's gold foil experiment showed that most of an atom's mass and positive charge is concentrated in a small, dense nucleus, with electrons orbiting around it.
- Example:** The nucleus of gold (Au) was shown to deflect alpha particles.

Bohr's Planetary Model

- Niels Bohr proposed that electrons move in fixed orbits around the nucleus, with energy levels. Electrons can jump between levels by absorbing or releasing energy.
- Example:** Hydrogen emits light when its electron transitions between energy levels.

Quantum Mechanical Model

- The modern model describes electrons as existing in probability clouds called orbitals rather than fixed paths.
- Example:** In a hydrogen atom, the electron cloud is spherical around the nucleus.

2. What are isotopes, and why are they important in science and technology?

Definition of Isotopes

Isotopes are atoms of the same element that have the same number of protons but different numbers of neutrons. This causes them to have identical chemical properties but different physical properties, such as mass and stability.

- Example:** Carbon-12 and Carbon-14 are isotopes of carbon.

Properties of Isotopes

- Isotopes have the same atomic number but different mass numbers.
- Some isotopes are stable (e.g., Carbon-12), while others are radioactive (e.g., Carbon-14).

Applications of Isotopes

1. Radiocarbon Dating:

- Carbon-14 is used to determine the age of archaeological artifacts and fossils.
- **Example:** The age of ancient wooden tools can be estimated by measuring the decay of Carbon-14.

2. Medical Applications:

- Radioisotopes like Cobalt-60 are used in cancer treatment to target and destroy cancerous cells.
- **Example:** Technetium-99m is used in imaging scans to diagnose organ issues.

3. Nuclear Energy:

- Uranium-235 is an isotope used as fuel in nuclear reactors to produce energy.
- **Example:** The fission of Uranium-235 releases energy and neutrons for chain reactions.

4. Scientific Research:

- Stable isotopes are used as tracers in chemical and biological studies.
- **Example:** Oxygen isotopes (O-16 and O-18) are used to study climate patterns.

3. How are cations and anions formed, and what roles do they play in chemistry?

Formation of Cations and Anions

1. Cations:

- Formed when an atom loses one or more electrons, resulting in a positive charge.
- **Example:** Sodium (Na) loses one electron to form Na^+ .

2. Anions:

- Formed when an atom gains one or more electrons, resulting in a negative charge.
- **Example:** Chlorine (Cl) gains one electron to form Cl^- .

Roles of Cations and Anions

1. Electrolytes in Solutions:

- Cations and anions dissolve in water to conduct electricity.
- **Example:** In table salt (NaCl), Na^+ and Cl^- ions dissociate in water, making it conductive.

2. Chemical Bonding:

- Cations and anions form ionic bonds to create stable compounds.
- **Example:** Sodium chloride (NaCl) is formed by the attraction between Na^+ and Cl^- ions.

3. Biological Functions:

- Ions like potassium (K^+) and calcium (Ca^{2+}) are vital for nerve impulses and muscle contraction.
- **Example:** Calcium ions (Ca^{2+}) are essential for bone formation and blood clotting.

4. Industrial Applications:

- Electroplating, battery production, and water purification rely on ions.
- **Example:** Anions like sulfate (SO_4^{2-}) are used in detergents and fertilizers.

4. What is electronic configuration, and how does it determine the chemical properties of elements?

Definition of Electronic Configuration

Electronic configuration describes how electrons are distributed in an atom's orbitals, based on their energy levels. It follows the **Aufbau principle** (fill lower energy orbitals first), **Pauli exclusion principle**, and **Hund's rule**.

- **Example:** Carbon's configuration is $1s^2 2s^2 2p^2$.

Determination of Chemical Properties

1. Valence Electrons:

- The number of electrons in the outermost shell (valence shell) determines an element's reactivity.
- **Example:** Sodium ($1s^2 2s^2 2p^6 3s^1$) has one valence electron, making it highly reactive.

2. Bond Formation:

- Elements with incomplete valence shells form bonds to achieve stability (octet rule).
- **Example:** Oxygen ($1s^2 2s^2 2p^4$) needs two more electrons to complete its octet, forming bonds with two hydrogen atoms in water (H_2O).

3. Periodic Trends:

- Elements with similar configurations have similar properties and belong to the same group in the periodic table.
- **Example:** Halogens (F, Cl, Br) have seven valence electrons, making them highly reactive nonmetals.

Applications of Electronic Configuration

1. Predicting Reactivity:

- Highly reactive elements like alkali metals and halogens can be identified by their configurations.
- **Example:** Fluorine ($1s^2 2s^2 2p^5$) readily gains one electron to form F^- .

2. **Spectroscopy:**

- Excited electrons release light of specific wavelengths, which is used to identify elements.
- **Example:** Sodium's yellow flame test corresponds to its electronic transitions.

3. **Chemical Bonding and Compounds:**

- Understanding configurations helps in predicting the types of bonds elements will form.
- **Example:** Carbon forms four covalent bonds due to its configuration ($1s^2 2s^2 2p^2$).

MCQS

1. **Who proposed the atomic model where atoms are indivisible particles?**
 - a) Rutherford
 - b) Dalton** ☒
 - c) Thomson
 - d) Bohr
2. **What is the main feature of Rutherford's atomic model?**
 - a) Atoms are indivisible
 - b) Atoms have a dense nucleus with electrons orbiting around it** ☒
 - c) Atoms are made of positive spheres
 - d) Atoms are cloud-like structures
3. **What did Bohr's atomic model introduce?**
 - a) Neutrons in the nucleus
 - b) Electrons in fixed orbits around the nucleus** ☒
 - c) Electrons as waves
 - d) Protons outside the nucleus
4. **What is the most accurate atomic model today?**
 - a) Thomson's model
 - b) Bohr's model
 - c) Quantum mechanical model** ☒
 - d) Rutherford's model
5. **Which subatomic particle was discovered first?**
 - a) Proton
 - b) Electron** ☒
 - c) Neutron
 - d) Quark
6. **What is the charge of a neutron?**
 - a) Positive
 - b) Negative
 - c) Neutral** ☒
 - d) Variable
7. **Which subatomic particle determines the identity of an element?**
 - a) Proton** ☒
 - b) Electron
 - c) Neutron
 - d) Nucleus
8. **Who discovered the proton?**
 - a) J.J. Thomson
 - b) James Chadwick
 - c) Ernest Rutherford** ☒
 - d) Niels Bohr
9. **What is the atomic number of an element?**
 - a) The number of protons in the nucleus** ☒

- b) The number of neutrons in the nucleus
c) The total number of subatomic particles
d) The atomic mass
10. Which of these isotopes is radioactive?
a) Carbon-12
b) Oxygen-16
c) Carbon-14 ☒
d) Hydrogen-1
11. What are isotopes?
a) Atoms of different elements with the same atomic number
b) Atoms of the same element with different numbers of neutrons ☒
c) Atoms with no electrons
d) Molecules with varying bonds
12. What is the relative atomic mass of an element?
a) The mass of an atom compared to carbon-14
b) The weighted average mass of its isotopes compared to carbon-12 ☒
c) The total protons and neutrons in the nucleus
d) The total electrons in the atom
13. What is the atomic mass unit (amu)?
a) The mass of one electron
b) 1/12th the mass of a carbon-12 atom ☒
c) The mass of one neutron
d) The mass of one hydrogen atom
14. What is the electronic configuration of carbon?
a) $1s^2 2s^2 2p^3$
b) $1s^2 2s^2 2p^2$ ☒
c) $1s^2 2p^6$
d) $1s^2 2s^1$
15. How many valence electrons does oxygen have?
a) 4
b) 6 ☒
c) 8
d) 2
16. What is the electronic configuration of sodium (Na)?
a) $1s^2 2s^2 2p^6$
b) $1s^2 2s^2 2p^6 3s^1$ ☒
c) $1s^2 2s^2 2p^5 3s^1$
d) $1s^2 2s^2$
17. What is the charge of a cation?
a) Positive ☒
b) Negative
c) Neutral
d) Variable
18. How is an anion formed?
a) By losing protons
b) By gaining electrons ☒

- c) By losing electrons
d) By gaining neutrons
19. Which is a common cation?
a) Na^+ ☒
b) Cl^-
c) SO_4^{2-}
d) HCO_3^-
20. Which is a common anion?
a) Mg^{2+}
b) K^+
c) Cl^- ☒
d) NH_4^+
21. What is the valency of chlorine?
a) 1 ☒
b) 2
c) 3
d) 4
22. Which principle explains the filling of electrons in orbitals?
a) Hund's rule
b) Aufbau principle ☒
c) Heisenberg uncertainty principle
d) Pauli exclusion principle
23. What does Hund's rule state?
a) Electrons fill the lowest energy level first
b) No two electrons can have the same quantum numbers
c) Electrons fill orbitals singly before pairing ☒
d) Electrons move in fixed paths
24. What is the electronic configuration of neon (Ne)?
a) $1s^2 2s^2 2p^3$
b) $1s^2 2s^2 2p^6$ ☒
c) $1s^2 2p^6$
d) $1s^2 2p^4$
25. What does the atomic number determine?
a) The number of protons ☒
b) The number of neutrons
c) The atomic mass
d) The number of electrons in isotopes
26. What charge does a sodium ion (Na^+) have?
a) +1 ☒
b) +2
c) -1
d) 0
27. What do isotopes of an element have in common?
a) The number of neutrons
b) The number of protons ☒
c) The mass number
d) The relative atomic mass

28. What property of isotopes makes them useful in radiocarbon dating?

- a) Stability
- b) Radioactivity ☒
- c) Large mass
- d) High reactivity

29. Which isotope is used in medical imaging?

- a) Carbon-12
- b) Cobalt-60
- c) Technetium-99m ☒
- d) Hydrogen-3

30. What does the relative atomic mass account for?

- a) The number of protons and neutrons
- b) The average mass of isotopes weighted by abundance ☒
- c) The atomic number
- d) The number of neutrons