Structure of Atom and Periodic classification of elements Main Contents

- ***Atomic Theory***
- ***Sub atomic particles and nucleus***
- ***Composition of atom and isotopes***
- ***Atomic models***
- ***Electromagnetic radiation and atomic spectra***
- ***Quantum mechanical model***
- ***Electron configuration***
- ***Traditional and modern periodic tables***
- ***Classification of elements***
- ***Periodic properties of elements***

🕮 Atomic Theory

Two thousand years ago, ancient Greek philosophers developed a theory of matter that was not based on the experimental evidences. A notable Greek philosopher, namely, **Democritus**, believed that all matter was composed of very tiny, indivisible particles. He called them **atomos**. Hence, the word 'atom' came from the Greek word atomos, which means uncuttable or indivisible.
Aristotle was part of the generation that succeeded **Democritus**. He did not believe in atomos. Aristotle thought that all matter was continuous. That is, if one proceeded on breaking down a substance, it would be impossible to reach to the last indivisible particle. In other words, it would continue to divide infinitely. His opinion was accepted for nearly 200 years.

- A theory of chemical combination, first stated by **John Dalton** in 1803. It involves the following postulates:
  i. *Elements consist of indivisible small particles (atoms).*
  ii. *All atoms of the same element are identical; different elements have different types of atom.*
  iii. *Atoms can neither be created nor destroyed.*

In 1804, John Dalton developed the first modern theory of atoms and proposed their existence. Dalton's atomic theory was based directly on the ideas of elements and compounds, and on the three laws of chemical combination. The three laws are:

i. The *law of conservation of mass* states that matter is neither created nor destroyed. This law is also called the law of indestructibility of matter. It means that the mass of the reactants is exactly equal to the mass of the products in any chemical reaction. A chemical reaction involves only the separation and union of atoms.
ii. The *law of definite proportions* states that a pure compound is always composed of the same elements combined in a definite ratio by mass. For example, water (H_2O) is composed of hydrogen and oxygen only. These elements are always in the proportion of 11.19% hydrogen to 88.81% oxygen by mass and in the proportion 2 : 1 by volume.
iii. The *law of multiple proportions* states that when two different compounds are formed from the same elements, the masses of one of the elements in the two compounds, compared to a given mass of the other element, is in a small whole-number ratio. For example, carbon and oxygen form two compounds: carbon monoxide and carbon dioxide. Carbon monoxide contains 1.3321 g of oxygen for each 1 g of carbon, whereas carbon dioxide contains 2.6642 g of oxygen for each 1g of carbon. Hence, carbon dioxide contains twice the mass of oxygen as does carbon monoxide.

❖ **The basic postulates of Dalton's Atomic Theory are summarized as follows:**

1. All elements are made up of small particles called atoms.
2. Atoms are indivisible and indestructible.
3. All atoms of a given element are identical in mass and in all other properties.
4. Atoms are neither created nor destroyed in chemical reactions.
5. Compounds are formed when atoms of more than one element combine.
6. In a given compound, the relative numbers and types of atoms are constant.

❖ **The Modern Atomic Theory can be summarized as follows:**

1. Atoms are the smallest particles of all elements that can take part in a chemical reaction.
2. An atom is divisible. It can be subdivided into electrons, protons, and neutrons.
   An atom is also indestructible *i.e.*, atoms can neither be created nor destroyed during ordinary chemical reactions.
3. Atoms of the same element may not be identical in mass because of the existence of isotopes.
4. Atoms of the same elements have identical chemical properties.
5. Atoms of different elements have different chemical properties.
6. Atoms of two or more elements combine in simple whole-number ratios to form compounds.

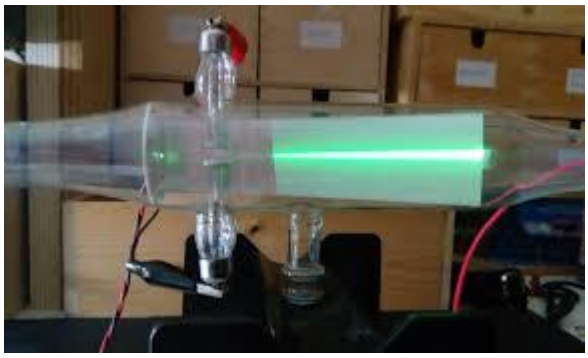
🕮 Sub atomic particles and nucleus

❖ **Discovery of electron**

In the late century, physicist J.J. Thomson began experimenting with *cathode ray tubes*. Cathode ray tubes are sealed glass a tube from which most of the air has been evacuated. A high voltage is applied across two electrodes at one end of the tube, which causes a beam of particles to flow from the cathode (the negatively-charged electrode) to the anode (the positively-charged electrode). The tubes are called cathode ray tubes because the particle beam or "cathode ray" originates at the cathode. The ray can be detected by painting a material known as *phosphors* onto the far end of the tube beyond the anode. The phosphors spark, or emit light, when impacted by the cathode ray.

Properties of cathode ray

Property 1: Cathode rays travel in a straight line and they can cast sharp shadows.
Property 2: They are negatively charged.
Property 3: Electric and magnetic fields deflect cathode rays.
Property 4: They are produced at the cathode and travel to the anode in a vacuum tube.



- J.J. Thomson's experiments with cathode ray tubes showed that *all* atoms contain tiny negatively charged subatomic particles or **electrons**.
- Thomson's **plum pudding model** of the atom had negatively-charged electrons

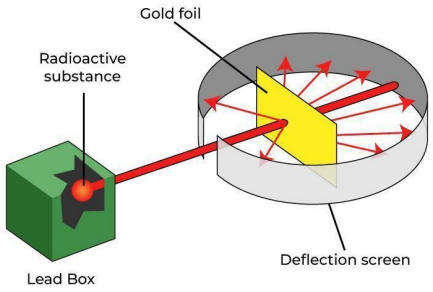
embedded within a positively-charged "soup."

- Rutherford's gold foil experiment showed that the atom is mostly empty space with a tiny, dense, positively-charged **nucleus**.
- Based on these results, Rutherford proposed the **nuclear model** of the atom.

Ernest Rutherford and the gold foil experiment

The discovery of the proton is credited to Ernest Rutherford, who proved that the nucleus of the hydrogen atom (i.e. a proton) is present in the nuclei of all other atoms in the year 1917.

- The next groundbreaking experiment in the history of the atom was performed by Ernest Rutherford, a physicist from New Zealand who spent most of his career in England and Canada. In his famous gold foil experiment, Rutherford fired a thin beam of particles at a very thin sheet of pure gold. Alpha particles are helium nuclei , and they are given off in various radioactive decay processes. In this case, Rutherford placed a sample of radium (a radioactive metal) inside a lead box with a small pinhole in it. Most of the radiation was absorbed by the lead, but a thin beam of particles escaped out of the pinhole in the direction of the gold foil. The gold foil was surrounded by a detector screen that would flash when hit with alpha particle.



Rutherford's Conclusion

1. Since most of the α-particles passed through the gold foil undeflected, most of the space in an atom is empty.
2. Some of the α-particles were deflected by small angles. This indicated the presence of a heavy positive centre in the atom, which Rutherford named the nucleus.
3. Only a few particles (1 in about a million) were either deflected by a very large angle or deflected back. This confirmed that the space occupied by the heavy positive centre must be very small.

	Mass		Charge	
	Grams	Atomic Mass units (u)	Coulombs	Electronic charge
Proton	1.67×10^{-24}	1.007276	$+ 1.602 \times 10^{-19}$	1+
Neutron	1.67×10^{-24}	1.008665	0	0
Electron	9.11×10^{-28}	0.0005486	-1.602×10^{-19}	1-

❖ **Discovery of Neutron**

The British physicist Sir James Chadwick discovered neutrons in the year 1932. He was awarded the Nobel Prize in Physics in the year 1935 for this discovery. It is important to note that the neutron was first theorized by Ernest Rutherford in the year 1920. Chadwick repeated their experiments but with the goal of looking for a neutral particle – one with the same mass as a proton, but with zero charge. His experiments were successful. “He was able to determine that the neutron did exist and that its mass was about 0.1 percent more than the proton's. “

❖ **Composition of atom and isotopes**

- The atomic number is simply the number of protons in the given element (Z), and the mass number is the number of protons + neutrons in an atom of the element (A).
- Isotopes are atoms of the same element that have the same number of protons (i.e., atomic number, "Z") but a different number of neutrons, meaning that their mass number, "A", varies. Take hydrogen, for example. It has three naturally occurring isotopes-- ^{1}H, ^{2}H, and ^{3}H.

Isotope	Nuclear symbol	No. of protons	No. of electrons	No. of neutrons	Z	A
Hydrogen-1 (protium)	$^{1}_{1}H$	1	1	0	1	1
Hydrogen-2 (deuterium)	$^{2}_{1}H$	1	1	1	1	2
Hydrogen-3 (tritium)	$^{3}_{1}H$	1	1	2	1	3
Helium-3	$^{3}_{2}He$	2	2	1	2	3
Helium-4	$^{4}_{2}He$	2	2	2	2	4

- Average atomic mass is the average weight of all the different isotopes of a particular element. And it depends on the percentage abundance of the isotope.
- Average atomic mass = f1M1+ f2M2 +… + fnMn where f is the fraction representing the natural abundance of the isotope and M is the mass number (weight) of the isotope.

Question 1 The average atomic mass of a sample of an element X is 16.2u. What are the percentages of isotopes 168X and 188X in the sample?

Solution: Isotope Percentage Calculation
Given:

- ✓ Average atomic mass of element X = 16.2 u
- ✓ Isotope 1: ^{16}X (mass = 16 u)
- ✓ Isotope 2: ^{18}X (mass = 18 u)

Step 1: Define variables
Let the percentage of ^{16}X = x%
Then the percentage of ^{18}X = (100 - x)%
Step 2: Use the average atomic mass formula
Average mass = $(mass_1 \times percentage_1 + mass_2 \times percentage_2) \div 100$
16.2 = [16 × x + 18 × (100 - x)] ÷ 100
Step 3: Solve for x
16.2 × 100 = 16x + 1800 - 18x
1620 = 1800 - 2x
2x = 1800 - 1620
2x = 180
x = 90
Step 4: Calculate both percentages

- ✓ ^{16}X = x% = **90%**

- ✓ ^{18}X = (100 - x)% = **10%**

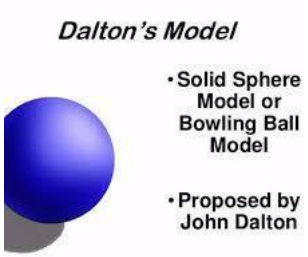
  Step 5: Verify

  (16 × 0.9) + (18 × 0.1) = 14.4 + 1.8 = 16.2 u ✓

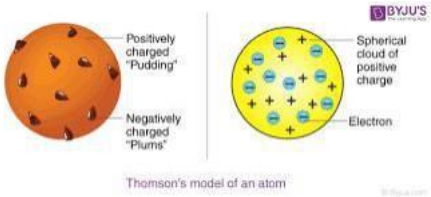
  Answer:
- ✓ ^{16}X = **90%**
- ✓ ^{18}X = **10%**

🕮 Atomic models

1, Dalton's Model of an Atom Based on all his observations, Dalton proposed his model of an atom. It is often referred to as the billiard ball model. He defined an atom to be a ball-like structure, as the concepts of atomic nucleus and electrons were unknown at the time.



2, According to the postulates of **Thomson's atomic model,** an atom resembles a sphere of positive charge with electrons (negatively charged particles) present inside the sphere. The positive and negative charge is equal in magnitude and therefore an atom has no charge as a whole and is electrically neutral. Thomson's model of the atom is also called the Plum pudding model, the Raisin pudding model and the Watermelon model.



3, **Rutherford's atomic model** became known as the nuclear model. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom.

4, Bohr Atomic model

Postulates of Bohr's Model of an Atom

- In an atom, electrons (negatively charged) revolve around the positively charged nucleus in a definite circular path called orbits or shells.
- Each orbit or shell has a fixed energy and these circular orbits are known as orbital shells.
- The energy levels are represented by an integer (n=1, 2, 3…) known as the quantum number. This range of quantum number starts from nucleus side with n=1 having the lowest energy level. The orbits n=1, 2, 3, 4… are assigned as K, L, M, N…. shells and when an electron attains the lowest energy level, it is said to be in the ground state.

- The electrons in an atom move from a lower energy level to a higher energy level by gaining the required energy and an electron moves from a higher energy level to lower energy level by losing energy.

Limitations of Bohr's Model of an Atom

- Bohr's model of an atom failed to explain the Zeeman Effect (effect of magnetic field on the spectra of atoms).
- It also failed to explain the Stark effect (effect of electric field on the spectra of atoms).
- It violates the hisnbergs uncertainty principle.
- It could not explain the spectra obtained from larger atoms.

5, The Quantum Mechanical Model of the Atom views electrons within an atom as waves, not as particles as previously believed. Because of this, electrons can be described as having a high probability of residing in specific orbitals, or regions of space, within the atom.

During the 1920's, the discovery of the wave-like properties of electrons led to the wave-mechanical model or quantum-mechanical model of atoms. In this model, the electrons are associated with definite energy levels, but their locations cannot be pinpointed. Instead, they are described in terms of the probability of being found in certain regions of space about the nucleus. These regions of space are called orbitals. An orbital is a particularly shaped volume of space where the probability of finding an electron is at a maximum.

🕮 *Electromagnetic radiation and atomic spectra*

Quantum theory and photon

In 1900, Max Planck, the German physicist, came to an entirely new view of matter and energy. He made a revolutionary proposal, **energy like matter is discontinuous**.

According to Planck, atoms and molecules could emit or absorb energy only in discrete quantities, like small packages or bundles. Each of these small "packets" of energy is called a **quantum**. The energy of a quantum is proportional to the frequency of the radiation. The energy *E* of a single quantum is given by: $E=h\nu$

Where *h* is called Planck's constant and *v* is the frequency of radiation. The value of Planck's constant is 6.63×10^{-34} J·s.

- **The photoelectric effect** is **the emission of electrons from a material caused by electromagnetic radiation** such as ultraviolet light.
- **The threshold frequency** is the lowest frequency below which the photoelectric effect does not occur. The threshold frequency is defined as the lowest incident radiation frequency below which photoelectric emission is absolutely impossible regardless of the strength of the incident radiation.
- **A photon** is a tiny particle that comprises waves of electromagnetic radiation. As shown by Maxwell, photons are just electric fields traveling through space. Photons have no charge, no resting mass, and travel at the speed of light.

The minimum energy required to remove an electron is:

$Eo=hv_0$

Where E_0 is the minimum energy (of the photon), and v_0 the threshold frequency.

A photon with energy less than *Eo* ($v<v0$) cannot remove an electron, or a light with a frequency less than the *v*0 produces no electrons. On the other hand, if a light has $v>v0$, the energy in excess of that required to remove the electron is given to the electron as kinetic energy (KE):

$KEe=1/2mv^2=hv-hv_0$

Where *KEe* is the kinetic energy of an electron, *m* is mass of an electron, *v* is the velocity of an electron, *hv* is the energy of an incident photon, and hv_0 is the energy required to remove an electron from the metal's surface.

$E=mc^2$

Rearranging this equation, we have

$m=E/c^2$

Where *E* is energy, *m* is mass, and *c* is speed of light.

The main significance of **Equation** is that **energy has mass**. Using this equation, we can calculate the mass associated with a given quantity of energy or the **apparent mass** of a photon. For electromagnetic radiation of wavelength λ, the energy of each photon is given by the expression:

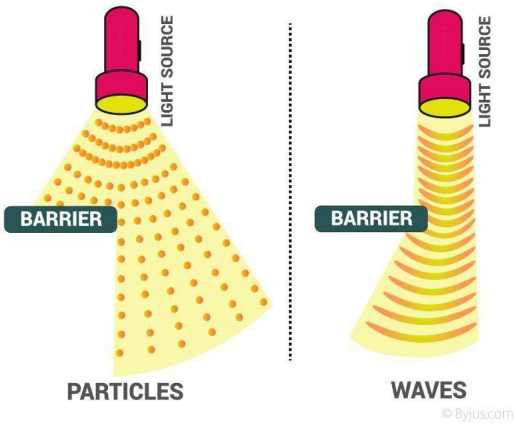
$Ephoton=hc\lambda$

Then, the apparent mass of a photon of light with wavelength is given by:

$m=E/c^2=hc/\lambda$

We can summarize the important conclusions from the work of Planck and Einstein as follows:

Energy is quantized. It can occur only in discrete units called **photon** or **quanta**. EMR, which was previously thought to exhibit only wave properties, seems to show certain characteristics of particulate matter as well. This phenomenon is sometimes referred to as the **dual nature of light** and is illustrated **Figure** .



Atomic or line spectra are produced from the emission of photons of electromagnetic radiation (light). Different kinds of spectrum are observed when an electric discharge, or spark, passes through a gas such as hydrogen. The electric discharge is an electric current that excites, or energizes, the atoms of the gas. More specifically, the electric current transfers energy to the electrons in the atoms raising them to excited states. The atoms then emit the absorbed energy in the form of light as the electrons return to lower energy states. When a narrow beam of this light is passed through a prism, we do not see a continuous spectrum, or rainbow, as sunlight does. Rather, only a few colors are observed, displayed as a series of individual lines. This series of lines is called the element's atomic spectrum or emission spectrum. The wavelengths of

these spectral lines are characteristic of the element producing them, and used for their identification. For example, the emission (line) spectrum of hydrogen atom.

- Changes in energy between discrete energy levels in hydrogen will produce only certain wavelengths of emitted light, For example, a given change in energy from a high to a lower level would give a wavelength of light which can be expressed using Planck's equation: $E = h\nu = hc/\lambda$ **The Quantum Mechanical Model** of the Atom views electrons within an atom as waves, not as particles as previously believed. Because of this, electrons can be described as having a high probability of residing in specific orbitals, or regions of space, within the atom.

- The Bohr model of an electron orbiting around the nucleus, looking like a planet around the Sun, doesn't explain properties of atoms. The planetary view of one charged particle orbiting another particle of opposite charge does not match some of the best-known laws of classical physics.

- In 1827, Werner Heisenberg, the German physicist, proposed the **Heisenberg uncertainty principle**, stating that it is not possible to know with great certainty both an electron's position and its momentum p*p* (where $p=mv$) at the same time. Heisenberg was able to show that if Δp is the uncertainty in the momentum and Δx is the uncertainty of the position of the electron, then: $(\Delta x)(\Delta p) \geq h4\pi$ or $(\Delta x)\,\Delta(mv) \geq h4\pi$
- This equation shows that if we measure the momentum of a particle precisely (i.e., if Δp is very small), then the position will be correspondingly less precise (i.e, Δx will become larger). In simpler terms, if we know precisely where a particle is, we cannot know precisely where it has come from or where it is going. If we know precisely how a particle is moving, we cannot know precisely where it is.
- The Heisenberg uncertainty principle has no practical consequence in everyday objects in the macroscopic world. However, it has very significant consequences when applied to atomic and subatomic particles.
- The quantum mechanical description of the atom is based on the assumption that there are waves associated with both matter and radiations. The mathematical description of such wave provides information about the energy of electrons and their position.

Practice Questions

1. Which of the following is NOT a postulate of Dalton's original atomic theory?
A) Elements consist of indivisible small particles called atoms.
B) All atoms of the same element are identical in mass and properties.
C) Atoms can be subdivided into electrons, protons, and neutrons.
D) Atoms of different elements combine in simple whole-number ratios to form compounds.

2. The law of definite proportions states that:
A) Mass is neither created nor destroyed in a chemical reaction.
B) A pure compound always contains the same elements in the same proportion by mass.
C) When two elements form different compounds, the masses of one element combining with a fixed mass of the other are in a small whole-number ratio.
D) Atoms of the same element may have different masses.

3. Which ancient Greek philosopher believed that matter was continuous and could be divided infinitely?
A) Democritus
B) Aristotle
C) Plato
D) Epicurus

4. The law of multiple proportions is best illustrated by:
A) H_2O and H_2O_2
B) NaCl and KCl
C) CO_2 and SO_2
D) CH_4 and C_2H_6

5. According to the modern atomic theory, atoms of the same element:
A) Must have identical masses.
B) Have identical chemical properties but may differ in mass.
C) Cannot be created or destroyed in nuclear reactions.
D) Are always electrically charged.

6. J.J. Thomson's cathode ray tube experiments led to the discovery of:
A) Proton
B) Neutron
C) Electron
D) Nucleus

7. In Rutherford's gold foil experiment, the observation that most alpha particles passed straight through the foil led to the conclusion that:
A) The nucleus is positively charged.
B) The atom is mostly empty space.
C) Electrons orbit the nucleus.
D) The nucleus is very small and dense.

8. Who discovered the neutron?
A) Ernest Rutherford
B) J.J. Thomson
C) James Chadwick
D) Niels Bohr

9. Which subatomic particle has a relative mass of approximately 1 and no charge?
A) Proton
B) Electron
C) Neutron
D) Positron

10. The atomic number (Z) of an element represents the number of:
A) Neutrons

B) Protons + neutrons
C) Protons
D) Electrons in a neutral atom

11. An isotope of an element has the same number of _____ but a different number of _____.
A) protons, electrons
B) neutrons, protons
C) protons, neutrons
D) electrons, neutrons

12. A sample of element X has an average atomic mass of 20.2 u. It consists of two isotopes: ^{20}X (mass 20 u) and ^{22}X (mass 22 u). The percentage abundance of ^{22}X is:
A) 10%
B) 20%
C) 80%
D) 90%

13. Which statement about isotopes is correct?
A) They have the same chemical properties but different physical properties.
B) They have different positions in the periodic table.
C) They have the same number of neutrons.
D) They are only found for radioactive elements.

14. The average atomic mass of chlorine is 35.5 u. Which pair of isotopes is most likely?
A) ^{35}Cl and ^{36}Cl
B) ^{35}Cl and ^{37}Cl
C) ^{34}Cl and ^{36}Cl
D) ^{35}Cl and ^{38}Cl

15. Thomson's model of the atom is often called the "plum pudding" model because:
A) Electrons are embedded in a positively charged sphere.
B) The nucleus is like a plum in a pudding.
C) Electrons orbit the nucleus like planets.
D) The atom is mostly empty space.

16. Which of the following is a limitation of Bohr's model of the atom?
A) It explains the spectra of hydrogen but not multi-electron atoms.
B) It violates the Heisenberg uncertainty principle.
C) It fails to explain the Zeeman and Stark effects.
D) All of the above.

17. In Rutherford's nuclear model, the nucleus contains:
A) Only protons
B) Only neutrons
C) Protons and neutrons
D) Protons and electrons

18. According to the quantum mechanical model, an orbital is:
A) A fixed circular path of an electron.
B) A region of space where the probability of finding an electron is high.
C) The same as an orbit in Bohr's model.
D) A path that electrons follow around the nucleus.

19. Which experiment provided evidence for the existence of the atomic nucleus?
A) Cathode ray tube experiment
B) Gold foil experiment
C) Photoelectric effect experiment
D) Oil drop experiment

20. The energy of a photon is given by $E=hv$. What does h represent?
A) Speed of light
B) Planck's constant
C) Threshold frequency
D) Wavelength

21. The photoelectric effect demonstrates that:
A) Light behaves only as a wave.
B) Light behaves only as a particle.
C) Light has dual nature (wave and particle).
D) Electrons are waves.

22. If the threshold frequency for a metal is v_0, what happens when light of frequency $v<v_0$ is incident?
A) Electrons are emitted with kinetic energy $h(v-v_0)$.
B) No electrons are emitted regardless of intensity.
C) Electrons are emitted only if intensity is high.
D) Electrons are emitted with zero kinetic energy.

23. The line (atomic) emission spectrum of hydrogen is produced when:
A) Electrons absorb energy and jump to higher levels.
B) Electrons return from excited states to lower energy levels, emitting photons.
C) The atom is in its ground state.
D) White light passes through hydrogen gas.

24. The Heisenberg uncertainty principle states that:
A) Energy and time cannot both be known precisely.
B) Position and momentum cannot both be known precisely.
C) Frequency and wavelength cannot both be known precisely.
D) Mass and velocity cannot both be known precisely.

25. Which quantum number describes the shape of an orbital?
A) Principal quantum number (n)
B) Azimuthal quantum number (l)

C) Magnetic quantum number (m_l)
D) Spin quantum number (m_s)

26. The maximum number of electrons that can occupy the third shell (n=3) is:
A) 8
B) 18
C) 32
D) 2

27. Which of the following electron configurations violates Hund's rule?
A) $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$
B) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1$
C) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^1$
D) $1s^2\ 2s^1\ 2p_x^1$

28. The modern periodic law states that the properties of elements are a periodic function of their:
A) Atomic mass
B) Atomic number
C) Atomic volume
D) Atomic radius

29. Which of the following elements would have the largest atomic radius?
A) Lithium (Li)
B) Beryllium (Be)
C) Boron (B)
D) Carbon (C)

30. Which group in the periodic table contains the most reactive nonmetals?
A) Group 1 (alkali metals)
B) Group 2 (alkaline earth metals)
C) Group 17 (halogens)
D) Group 18 (noble gases)

ANSWER KEY

Q#	Answer
1	C
2	B
3	B

Q#	Answer
4	A
5	B
6	C
7	B
8	C
9	C
10	C
11	C
12	A
13	A
14	B
15	A
16	D
17	C
18	B
19	B
20	B
21	C
22	B
23	B

Q#	Answer
24	B
25	B
26	B
27	B
28	B
29	A
30	C