



APPLIED CHEMISTRY-I

For First Semester Diploma Students



Dr. Indu Singh

As per AICTE and UBTER, Uttarakhand Syllabus

KHANNA PUBLISHERS

Investing in Learning

Applied Chemistry – I

*As per UBTER, Uttarakhand and
AICTE Syllabus*

for

First Semester Diploma Students

*by
Dr. Indu Singh*



KHANNA PUBLISHERS

Operational Office:

4575/15, Onkar House, Ground Floor,
Darya Ganj, New Delhi 110 002

Phones : 011-45033819 • Mob. 09811541460

email : contactus@khannapublishers.in

Published by :

Romesh Chander Khanna & Vineet Khanna
for **KHANNA PUBLISHERS**
2-B, Nath Market, Nai Sarak,
Delhi-110006 (India)

Visit us at : www.khannapublishers.in

Copyright: Publishers

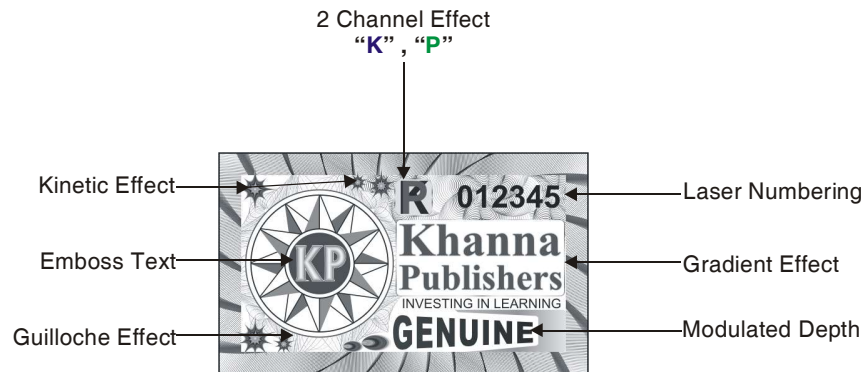
© 1979 and onward.

This book or part thereof cannot be translated or reproduced in any form without the written permission of Publisher.

Hologram & Description

To all readers of our books, to prevent yourself from being frauded by pirated books. Please make sure that there is an Hologram on the cover of our books with the below specifications. If you find any book without the Hologram or Description, please mail us at contactus@khannapublishers.in.

Thanking you



ISBN No. : 978-93-87394-64-3

First Edition : 2019

Preface

The book Applied Chemistry—I, has been written for the Diploma students of UBTER, Uttarakhand as per AICTE curriculum for their first semester diploma course. The aim of writing this book is to be provided the students a clear understanding of the basic concepts and principles.

The subject matter has been developed in steps for easy understanding. Throughout the book, the stress has been given on the topics through figures and solved examples.

The book is divided into **Eight Chapters**. At the end of each chapter, a brief summary is given for readers reference.

Exercises, Problems and Multiple Choice Questions (MCQs) with their answers and hints are provided at the end of each unit.

In spite of our efforts, it is possible that some errors might have crept in. We shall acknowledge with gratitude, if any such error is brought in our notice.

Also, any suggestions and comments from students and teachers for improvement of the book are welcome.

We hope the book will be found useful by the readers.

—*Publishers*

Contents

Chap.		Pages
1.	Atomic Structure	1–19
1.1.	Introduction	1
1.2.	Fundamental Particles <i>i.e.</i> , Proton, Neutron and Electron	2
1.3.	Electronic Configuration of Elements (upto $z = 30$)	3
1.4.	Atomic Mass, Molecular Mass and Equivalent Mass	8
	<i>Summary</i>	12
	<i>Solved Examples</i>	13
	<i>Exercise</i>	15
	<i>Problems</i>	15
	<i>Multiple Choice Questions</i>	17
2.	Chemical Bonding	20–40
2.1.	Introduction	20
2.2.	Hydrogen Bonding	30
	<i>Summary</i>	36
	<i>Solved Examples</i>	36
	<i>Exercise</i>	37
	<i>Problems</i>	38
	<i>Multiple Choice Questions</i>	38
3.	Classification of Elements	41–70
3.1.	Introduction	41
3.2.	Modern Periodic Law	41
3.3.	Study of Periodicity in Physical and Chemical Properties	49
3.4.	Variation of Effective Nuclear Charge in a Period and Metallic Characters	62
	<i>Summary</i>	64
	<i>Solved Examples</i>	64
	<i>Exercise</i>	65
	<i>Problems</i>	66
	<i>Multiple Choice Questions</i>	68
4.	Water Treatment	71–108
4.1.	Introduction	71
4.2.	Sources of Water	71
4.3.	Softening of Water	80
4.4.	Internal Treatment of Water	85
4.5.	Determination of Hardness of Water	87
4.6.	Qualities of Water Used for Drinking Purposes	91
	<i>Summary</i>	100
	<i>Solved Examples</i>	101

<i>Chap.</i>	<i>Pages</i>
<i>Exercise</i>	... 105
<i>Problems</i>	... 106
<i>Multiple Choice Questions</i>	... 107
5. Solutions	... 109–132
5.1. Introduction	... 109
5.2. Concept of Solution	... 109
5.3. Strength of a Solution	... 113
5.4. Buffer Solutions	... 115
5.5. Definition of pH Value	... 120
<i>Summary</i>	... 124
<i>Solved Examples</i>	... 124
<i>Exercise</i>	... 128
<i>Problems</i>	... 128
<i>Multiple Choice Questions</i>	... 130
6. Electrochemistry	... 133–152
6.1 Definition of the Electrochemistry	... 133
6.2. Faraday's Laws of Electrolysis with Simple Numerical Problems	... 141
6.3. Different Industrial Applications of Electrolysis, with Special Reference to Electroplating, Electrefining and Electrometallurgy	... 144
<i>Summary</i>	... 148
<i>Solved Examples</i>	... 148
<i>Exercise</i>	... 150
<i>Problems</i>	... 151
<i>Multiple Choice Questions</i>	... 152
7. Catalyst	... 153–166
7.1. Definition of Catalyst	... 153
7.2. Characteristics of Catalytic Reaction	... 158
7.3. Theory of Catalysis and Mechanism of Catalysis	... 160
<i>Summary</i>	... 164
<i>Exercise</i>	... 165
<i>Multiple Choice Questions</i>	... 165
8. Environmental Chemistry	... 167–200
8.1. General Concept of Pollution	... 167
8.2. Factor Effecting Air, Water, Noise and Soil Pollution	... 170
8.3. Green House Effect, Ozone Layer Depletion, Global Warming, Acid Rain and Smog	... 185
<i>Summary</i>	... 196
<i>Exercise</i>	... 197
<i>Multiple Choice Questions</i>	... 198

1

Atomic Structure

1.1. INTRODUCTION

Chemistry has been defined as the study of matter in terms of its structure, composition and the properties. As you are aware, matter is made up of atoms, and therefore an understanding of the structure of atom is very important. You have studied in your earlier classes that the earliest concept of atom (smallest indivisible part of matter) was given by ancient (600-400 BC) Indian and Greek philosophers. At that time, there were no experimental evidence. The origin of the concept of atom was based on their thoughts on 'What would happen, if we continuously keep dividing matter'. John Dalton revived the concept of atom in the beginning of nineteenth century in terms of his atomic theory which successfully explained the laws of chemical combination. Later experiments showed that the atom is not indivisible but has an internal structure.

An atom is the smallest constituent unit of ordinary matter that has the properties of a chemical element. Every solid, liquid, gas, and plasma is composed of neutral or ionized atoms. Atoms are very small; typical sizes are around 100 picometers (a ten-billionth of metre, in the short scale).

Atoms are small enough that attempting to predict their behavior using classical physics as if they were billiard balls, for example gives noticeable incorrect predictions due to quantum effects. Through the development of physics, atomic models have incorporated quantum principles to better explain and predict the behaviour.

Every atom is composed of a nucleus and one or more electrons bound to the nucleus. The nucleus is made of one or more protons and typically a similar number of neutrons. Protons and neutrons are called nucleons. More than 99.94% of an atom's mass is in the nucleus. The protons have a positive electric charge, the electrons have a negative electric charge and the neutrons have no electric charge. If the number of protons and electrons are equal, that atom is electrically neutral. If an atom has more or fewer electrons than protons, then it has an overall negative or positive charge, respectively, and it is called an ion.

A matter is anything that has mass and occupies space. Pen, paper, clips, sand, air, ice, etc. are different forms of matter. Every matter is made up of tiny particles. These particles are so tiny that they can't be seen with naked eyes.

1.2. FUNDAMENTAL PARTICLES i.e., PROTON, NEUTRON AND ELECTRON

A fundamental particle—also known as an elementary particle—is one that cannot be broken into smaller pieces. An atom is not a fundamental particle, because it consists of protons and neutrons in its nucleus and has a cloud of electrons orbiting the nucleus. Protons and neutrons are also not fundamental particles, because they consist of quarks. However, quarks and electrons cannot be broken apart and are fundamental particles. (Fig. 1.1)

In 1897, J.J. Thomson discovered electron as a constituent of atom. He determined that electrons had a negative charge and had very little mass as compared to that of the atom. Since an atom was found to be electrically neutral it was inferred that some source of positive charge must be present in the atom. This soon led to the experimental discovery of the proton, which is a positively charged sub-atomic particle. The existence of protons in the atoms was discovered by E. Goldstein. Proton was found approximately 1840 times heavier than an electron. Further experiments revealed that the atomic masses were more than that expected from the presence of just protons and electrons in the atom. For example, the mass of helium atom was expected to be double that of hydrogen atom but was actually found to be almost four times the mass of hydrogen atom. This suggested the presence of neutral particles with mass comparable to that of protons in the atom. Sir James Chadwick discovered this neutral particle and called it neutron subsequently in 1932.

1.2.1. Proton

The protons are positively charged particles found in the atoms of all the elements. They are located in the nucleus. The proton is usually represented by symbol p^+ .

1.2.2. Neutron

The neutron is a neutral particle found in the nucleus. It is represented by n .

1.2.3. Electron

The electrons are negatively charged particles found in the atoms of all the elements. They are located outside of nucleus. An electron is represented by symbol e^- .

The atoms are not individual but are made up of three fundamental particles whose mass, charge, and location are given in Table 1.1.

Table 1.1. Mass of fundamental particles of atom and locating of charge

Particle	Symbol	Mass / kg	Actual Charge / C	Relative charge
Electron	e	$9.109\ 389 \times 10^{-31}$	$-1.602\ 177 \times 10^{-19}$	-1
Proton	p	$1.672\ 623 \times 10^{-27}$	$1.602\ 177 \times 10^{-19}$	+1
Neutron	n	$1.674\ 928 \times 10^{-27}$	0	0

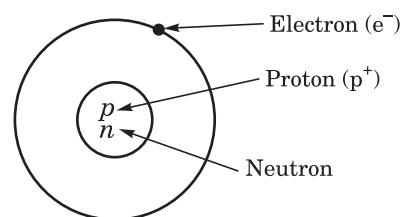


Fig. 1.1

1.2.4. Characteristics of Fundamental Particles

(i) Characteristics of Electrons

- **Charge** : Electron is a negatively charged particle.
- **Magnitude of charge** : Charge of electron is 1.6022×10^{-19} coulomb.
- **Mass of electron** : Mass of electron is 0.000548597 a.m.u. or 9.11×10^{-31} kg.

- *Symbol of electron* : Electron is represented by “e”.
- *Location in the atom* : Electrons revolve around the nucleus of atom in different circular orbits.

(ii) **Characteristics of Protons**

- *Charge* : Proton is a positively charged particle.
- *Magnitude of charge* : Charge of proton is 1.6022×10^{-19} coulomb.
- *Mass of proton* : Mass of proton is 1.0072766 a.m.u. or 1.6726×10^{-27} kg.
- *Comparative mass*: Proton is 1837 times heavier than an electron.
- *Position in atom* : Protons are present in the nucleus of atom.

(iii) **Characteristics of Neutrons**

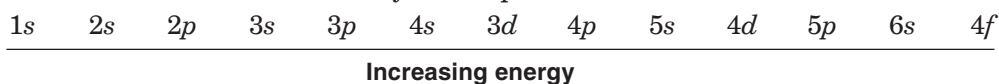
- *Charge* : Neutron is a neutral particle because it has no charge.
- *Mass of neutron* : Mass of neutron is 1.0086654 a.m.u. or 1.6749×10^{-27} kg.
- *Comparative mass* : Neutron is 1842 times heavier than an electron.
- *Location in the atom* : Neutrons are present in the nucleus of an atom.

1.3. ELECTRONIC CONFIGURATION OF ELEMENTS (UPTO Z = 30)

The distribution of electrons of the atom in different orbitals is known as the **ground state electronic configuration of the atom**. The term **ground state** represents the position of electrons in the **lowest energy state** or the **most stable state**. The filling of orbitals in an atom is governed by the following three fundamental rules.

1.3.1. Aufbau's Rule

According to this rule “**the orbitals are filled progressively in order of their increasing energy beginning with the orbital of lowest energy**”. The order of increasing energies of various orbitals based on $(n + l)$ rule may be compiled as :



This trend of increasing energies of various orbitals may be remembered in the form of diagram represented in Fig. 1.2.

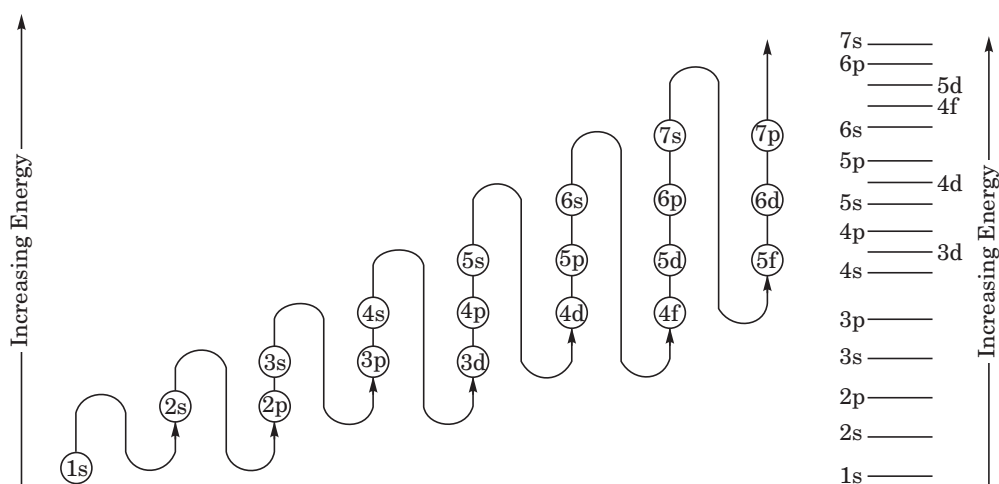


Fig. 1.2. Increasing order of energy of various orbitals

1.3.2. Pauli's Exclusion Principle

According to this principle **“no two electrons in an atom can have the same set of all the four quantum numbers”**. To illustrate this concept let us take an orbital specified by three quantum numbers n , l and m . Since an electron in an orbital can have two values of 's' as $+1/2$ or $-1/2$, therefore, the two electrons have different set of four quantum numbers. For example,

$2s$ orbital has $n = 2$, $l = 0$, $m = 0$ and $s = +\frac{1}{2}$ or $-\frac{1}{2}$.

Hence, one $2s$ electron has quantum numbers as $n = 2$, $l = 0$, $m = 0$ and $s = +1/2$ and second electron has quantum numbers as $n = 2$, $l = 0$, $m = 0$ and $s = -1/2$

Thus, the Pauli's exclusion principle can be redefined as:

- (i) an orbital cannot have more than two electrons.
- (ii) if orbital has two electrons, these must have opposite spins *i.e.*, they exist as paired electrons.

From the above discussion, we can observe that

- (i) s -subshell (containing only **one orbital**) can have a maximum of **two electrons**.
- (ii) p -subshell (containing **three orbitals**) can have a maximum of **six electrons**.
- (iii) d -subshell (containing **five orbitals**) can have a maximum of **ten electrons**.
- (iv) f -subshell (containing **seven orbitals**) can have a maximum of **fourteen electrons**.

From this knowledge we can calculate the total number of electrons in various shells as shown in table 1.2.

Table 1.2

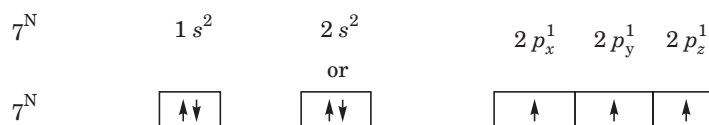
$(n = 1)$ K-shell	$(l = 0)$ 1s-subshell	$(m = 0)$ one orbital	2 electrons
$(n = 2)$ L-shell	$(l = 0)$ $2s$ -subshell	$(m = 0)$ one orbital	2 electrons
	$(l = 1)$ $2p$ -subshell	$(m = -1, 0, +1)$ <u>three orbitals</u> 4 orbitals	6 electrons <hr/> 8 electrons
$(n = 3)$ M-shell	$(l = 0)$ $3s$ -subshell	$(m = 0)$ one orbital	2 electrons
	$(l = 1)$ $3p$ -subshell	$(m = -1, 0, +1)$ three orbitals	6 electrons
	$(l = 2)$ $3d$ -subshell	$(m = -2, -1, 0, +1, +2)$ <u>five orbitals</u> 9 orbitals	10 electrons <hr/> 18 electrons
$(n = 4)$ N-shell	$(l = 0)$ $4s$ -subshell	$(m = 0)$ one orbital	2 electrons
	$(l = 1)$	$(m = -1, 0, +1)$	6 electrons

	4 <i>p</i> -subshell (<i>l</i> = 2)	three orbitals (<i>m</i> = - 2, - 1, 0, + 1, + 2)	10 electrons
	4 <i>d</i> -subshell (<i>l</i> = 3)	five orbitals (<i>m</i> = - 3, - 2, - 1, 0, + 1, + 2, + 3)	14 electrons
	4 <i>f</i> -subshell	seven orbitals	
		<u>16 orbitals</u>	<u>32 electrons</u>

1.3.3. Hund's Rule of Maximum Multiplicity

This rule governs the arrangement of electrons in the orbitals of identical energies *e.g.*, amongst p_x, p_y and p_z or amongst $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ and d_z^2 . According to this rule, electron pairing in orbitals of same energy (*i.e.* orbitals of *p*, *d* or *f* subshell) shall take place only when all the orbitals of the subshell contain one electron each.

This implies that two negatively charged electrons in a particular orbital experience greater repulsion and thus the pairing of electrons is avoided as long as it is possible during filling of orbitals of equal energy. Furthermore, the orbitals occupying single electron in parallel spin has higher stability because it corresponds to the state of lower energy. For example, electronic configuration of nitrogen (At. No. 7), applying Hund's rule, has three unpaired electrons in $2p$ orbitals.



Applying these rules, the electronic configuration of first thirty elements are summarized below in Tables 1.3 and 1.4. From sodium (At. No. 11) to argon (At. No. 18), $3s$ and $3p$ orbitals are successively filled. After $3p$, the 19th electron in potassium (At. No. 19) enters the $4s$ orbital instead of $3d$ because of Aufbau's rule.

Table 1.3. Electronic Configuration of the First 20 Elements

At. No	Element	Electronic configuration			
		K	L	M	N
1.	H	$1s^1$			
2.	He	$1s^2$			
3.	Li	$1s^2$	$2s^1$		
4.	Be	$1s^2$	$2s^2$		
5.	B	$1s^2$	$2s^2 2p_x^1$		
6.	C	$1s^2$	$2s^2 2p_x^1 2p_y^1$		
7.	N	$1s^2$	$2s^2 2p_x^1 2p_y^1 2p_z^1$		
8.	O	$1s^2$	$2s^2 2p_x^2 2p_y^1 2p_z^1$		
9.	F	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^1$		
10.	Ne	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$		
11.	Na	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^1$	
12.	Mg	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2$	

13.	Al	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^1$	
14.	Si	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^1 3p_y^1$	
15.	P	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^1 3p_y^1 3p_z^1$	
16.	S	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^2 3p_y^1 3p_z^1$	
17.	Cl	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^2 3p_y^2 3p_z^1$	
18.	Ar	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^2 3p_y^2 3p_z^2$	
19.	K	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^2 3p_y^2 3p_z^2$	$4s^1$
20.	Ca	$1s^2$	$2s^2 2p_x^2 2p_y^2 2p_z^2$	$3s^2 3p_x^2 3p_y^2 3p_z^2$	$4s^2$

The electronic configuration of next ten elements, known as transition elements, are given in Table 1.4. These elements (from scandium to zinc) have atomic numbers 21 to 30 and follow calcium. During filling of electrons in these elements, the electrons are added in the $3d$ orbitals.

Table 1.4: Electronic Configuration of Transition Elements

At. No.	Element	Electronic configuration
21.	Sc	$[\text{Ar}] 3d_{xy}^1$ $4s^2$
22.	Ti	$[\text{Ar}] 3d_{xy}^1 3d_{yz}^1$ $4s^2$
23.	V	$[\text{Ar}] 3d_{xy}^1 3d_{yz}^1 3d_{xz}^1$ $4s^2$
24.	Cr	$[\text{Ar}] 3d_{xy}^1 3d_{yz}^1 3d_{xz}^1 3d_{x^2-y^2}^1 3d_z^2$ $4s^1$
25.	Mn	$[\text{Ar}] 3d_{xy}^1 3d_{yz}^1 3d_{xz}^1 3d_{x^2-y^2}^1 3d_z^2$ $4s^2$
26.	Fe	$[\text{Ar}] 3d_{xy}^2 3d_{yz}^1 3d_{xz}^1 3d_{x^2-y^2}^1 3d_z^2$ $4s^2$
27.	Co	$[\text{Ar}] 3d_{xy}^2 3d_{yz}^2 3d_{xz}^1 3d_{x^2-y^2}^1 3d_z^2$ $4s^2$
28.	Ni	$[\text{Ar}] 3d_{xy}^2 3d_{yz}^2 3d_{xz}^2 3d_{x^2-y^2}^1 3d_z^2$ $4s^2$
29.	Cu	$[\text{Ar}] 3d_{xy}^2 3d_{yz}^2 3d_{xz}^2 3d_{x^2-y^2}^2 3d_z^2$ $4s^1$
30.	Zn	$[\text{Ar}] 3d_{xy}^2 3d_{yz}^2 3d_{xz}^2 3d_{x^2-y^2}^2 3d_z^2$ $4s^2$

Exceptional configuration of Cr and Cu

Elements with atomic numbers 24 and 29 have shown little deviation from Aufbau's rule. Here one electron from $4s$ orbital shifts to higher energy level $3d$ orbital. This migration of electron is due to higher stability of half-filled and fully-filled orbitals. The electronic configuration of these elements is given below.

Element	Atomic number	Expected electronic configuration	Observed electronic configuration
Cr	24	$[\text{Ar}]^{18} 3d^4 4s^2$	$[\text{Ar}]^{18} 3d^5 4s^1$
Cu	29	$[\text{Ar}]^{18} 3d^9 4s^2$	$[\text{Ar}]^{18} 3d^{10} 4s^1$

Stability of fully filled and half filled orbitals

The reasons for higher stability of Cr (half-filled orbitals) and Cu (fully-filled orbitals) can be explained as follows.

- (i) **Symmetry amongst orbitals** : Orbitals which are either half-filled or fully-filled are more symmetrical and therefore possess lower energy *i.e.*; extra stability. Two negatively charged electrons when present in one orbital will have higher electron-electron repulsion. Therefore, higher stability of unpaired electrons in chromium is because of less electron-electron repulsion.

- (ii) When orbitals are half-filled or fully-filled, the exchange of electrons between orbitals is maximum. Such exchange leads to greater stability of electrons in the orbitals, because low exchange energy results in higher stabilization energy.

Rules for determining the group number of element by its electronic configuration

- (i) If the last orbit (shell) has 1 or 2 electrons, then the group number of element is 1 or 2 respectively.
- (ii) If the last orbit (shell) has 3 or more electrons then the group number is the sum of electrons in the last shell plus 10.
- (iii) If electrons are present in $(n - 1)$ d orbital (*i.e.* penultimate shell) in addition to those in ns shell, then the group number is equal to the total number of electrons in ns and $(n - 1)$ d orbitals.

1.3.4. Heisenberg's Uncertainty Principle

One of the important consequences of the dual nature of an electron is the uncertainty principle given by Werner Heisenberg (1927). This principle is an important feature of wave mechanics and discusses the relationship between the position and momentum of a moving particle.

Heisenberg's uncertainty principle states that it is impossible to determine simultaneously the exact velocity (or momentum) and the exact position of a moving microscopic particle (*e.g.*, electron).

Accordingly, both momentum and position of a particle cannot be determined exactly at any instant. This implies that if momentum is measured very accurately, the measurement of position becomes less precise and *vice versa*.

If Δp (Δmv) is the uncertainty in measurement of momentum and Δx as the uncertainty in determination of position, then according to Heisenberg's principle the product of these two quantities is greater than or equal to $h/4\pi$, where h is Planck's constant

$$\Delta p \times \Delta x \geq \frac{h}{4\pi}$$

$$\Delta mv \times \Delta x \geq \frac{h}{4\pi}$$

$$\Delta v \times \Delta x \geq \frac{h}{4\pi m}$$

Explanatory proof of Heisenberg's uncertainty principle. A particle is visible only when it reflects or any other radiation from its surface. In order to locate the position of a revolving electron, one has to use bright light and super microscope. Bright light of very small wavelength or high energy when strikes an electron and the latter becomes visible, then energy of light is transferred to the electron. This results in change of momentum of electron. Thus the position of the electron though ascertained but its momentum changes drastically. However, if radiations of large wavelength or low energy is employed then the position of electron can't be determined with accuracy though its momentum is not altered. Therefore, it

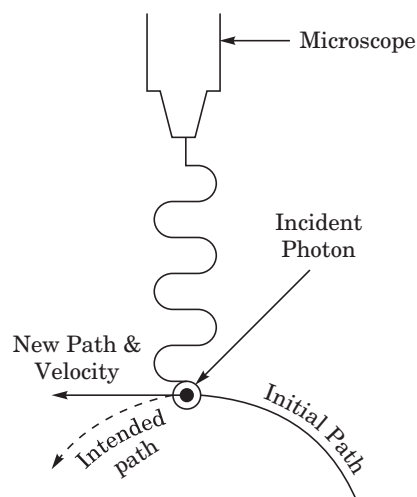


Fig. 1.3. The uncertainty principle

may be concluded that **it is not possible to determine simultaneously both the position and momentum of a small moving particle with absolute accuracy.**

1.4. ATOMIC MASS, MOLECULAR MASS AND EQUIVALENT MASS

We have already learn some basic idea about the terms atoms and molecules. Now, we shall learn about the atomic, molecular and equivalent masses.

1.4.1. Atomic Mass

The mass of an atom or the atomic mass is actually very-very small because atoms are extremely small. The mass spectrometry for determining the atomic masses fairly accurately. But, in the nineteenth century, scientists could determine mass of one atom **relative** to another by experimental means, as has been mentioned earlier. Hydrogen, being lightest atom was arbitrarily assigned a mass of 1 (without any units) and other elements were assigned masses relative to it. However, the present system of atomic mass is based on carbon-12. Here, Carbon-12 (C-12) is one of the isotopes of carbon and can be represented as ^{12}C . In this system, ^{12}C is assigned a mass of exactly 12 atomic mass unit (**amu**) and masses of all other atoms are given relative to this standard. One **atomic mass unit** is defined as a mass exactly equal to one twelfth the mass of one carbon-12 atom.

$$1 \text{ amu} = 1.66056 \times 10^{-24} \text{ g}$$

$$\text{Mass of an atom of hydrogen} = 1.6736 \times 10^{-24} \text{ g}$$

Thus, in terms of amu, the mass of hydrogen atom

$$\begin{aligned} &= \frac{1.6736 \times 10^{-24} \text{ g}}{1.66056 \times 10^{-24} \text{ g}} \\ &= 1.0078 \text{ amu} = 1.0080 \text{ amu} \end{aligned}$$

Similarly, the mass of oxygen-16 (^{16}O) atom would be 15.995 amu.

Today, '**amu**' has been replaced by '**u**' which is known as **unified mass**.

If we use atomic masses of elements in calculations, we actually use *average atomic masses* of elements which are explained below.

1.4.1.1. Average Atomic Mass

Some elements exist as more than one isotope if we take into account the existence of these isotopes and their relative abundance (per cent occurrence) the average atomic mass of that element can be computed. For example, carbon (C) has the following three isotopes with relative abundances and masses as shown against each of them.

Isotope	Relative Abundance (%)	Atomic Mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2×10^{-10}	14.00317

The average atomic mass of carbon will come out to be :

$$(0.98892) (12 \text{ u}) + (0.01108) (13.00335 \text{ u}) + (2 \times 10^{-10}) (14.00317 \text{ u}) = 12.011\text{u}$$

Similarly, the average atomic masses for other elements can be calculated. In the periodic table of elements, the atomic masses mentioned for different elements actually represented their average atomic masses.

Atomic Number

Atoms of each element contain a characteristic number of protons. In fact, the number of protons determines what atom we are looking at (e.g., all atoms with six protons are carbon atoms); the number of protons in an atom is called the atomic number. In contrast, the number of neutrons for a given element can vary. Forms of the same atoms that differ only in their number of neutrons are called isotopes. Together, the number of protons and the number of neutrons determine an element's mass number: mass number = protons + neutrons. If you want to calculate how many neutrons an atom has, you can simply subtract the number of protons, or atomic number, from the mass number.

Since an element's isotopes have different atomic masses, scientists may also determine the relative atomic mass—sometimes called the atomic weight, for an element. The relative atomic mass is an average of the atomic masses of all the different isotopes in a sample, with each isotope's contribution to the average determined by how big a fraction of the sample it makes up. The relative atomic masses given in periodic table entries—like the one for hydrogen, below—are calculated for all the naturally occurring isotopes of each element, weighted by the abundance of those isotopes on earth. Extraterrestrial objects, like asteroids or meteors, might have very different isotope abundances.

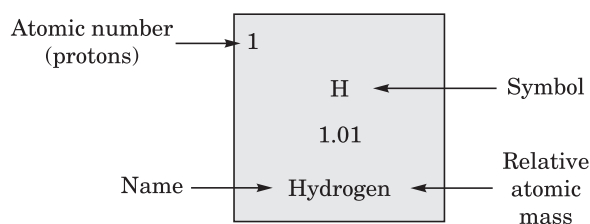


Fig. 1.4

What is an atom's atomic number?

The number of protons in the nucleus of an atom determines an element's atomic number. In other words, each element has a unique number that identifies how many protons are in one atom of that element. For example, all hydrogen atoms, and only hydrogen atoms, contain one proton and have an atomic number of 1. All carbon atoms, and only carbon atoms, contain six protons and have an atomic number of 6. Oxygen atoms contain 8 protons and have an atomic number of 8. The atomic number of an element never changes, meaning that the number of protons in the nucleus of every atom in an element is always the same.

What is an atom's mass number?

The mass number (symbol A), also called atomic mass number or nucleon number, is the total number of protons and neutrons (together known as nucleons) in an atomic nucleus.

The mass number is written either after the element name or as a superscript to the left of an element's symbol. For example, the most common isotope of carbon is carbon-12, or ^{12}C , which has 6 protons and 6 neutrons.

All atoms have a mass number which is derived as follows.

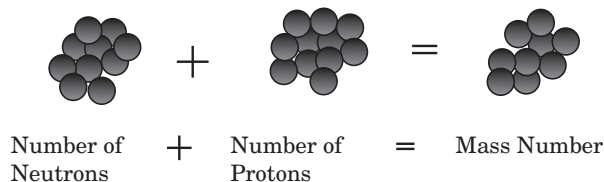


Fig. 1.5. Atomic mass number

Isotopes and isobars

Isotopes

Those elements that have the same atomic number but a different mass number are referred to as isotopes. Isotopes occur due to the presence of a different number of neutrons in elements having the same atomic number as mass number is the sum of the number of the neutrons and protons. Many but not all elements have isotopes. The isotopes of hydrogen are protium (has one proton and no neutrons), deuterium (has one proton and one neutron) and tritium (has one proton and two neutrons). The chemical properties of isotopes are same owing to the fact that they have the same number of protons and hence same number of electrons which determines the chemical properties of an element.

Isotopes of hydrogen

Hydrogen has three main isotopes; protium (${}^1\text{H}$), deuterium (${}^2\text{H}$) and tritium (${}^3\text{H}$). These isotopes form naturally in nature. Protium and deuterium are stable. Tritium is radioactive and has a half-life of about 12 years.

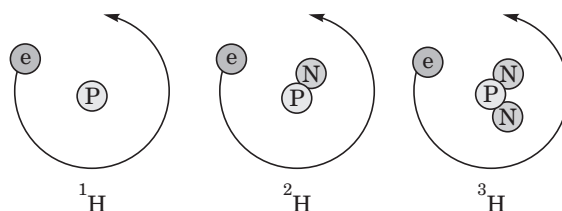


Fig. 1.6. Isotopes of hydrogen

Isotopes of oxygen

There are three stable isotopes of oxygen (${}_8\text{O}$): ${}^{16}\text{O}$, ${}^{17}\text{O}$, and ${}^{18}\text{O}$. Radioactive isotopes with mass number from ${}^{12}\text{O}$ to ${}^{24}\text{O}$ have also been characterized, all short-lived, with the longest-lived being ${}^{15}\text{O}$ with a half-life of 122.24 seconds. The shortest-lived is ${}^{12}\text{O}$ with a half-life of $580(30) \times 10^{-24}$ second.

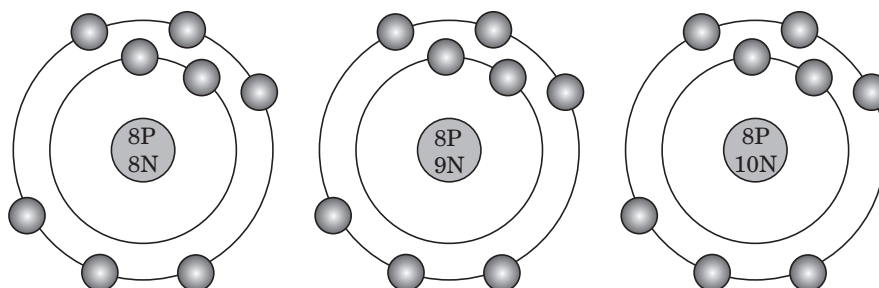


Fig. 1.7. Isotopes of oxygen